CLASS 9 SCIENCE

Chapter 8 Notes

Journey Inside the Atom

Chapter focus
This chapter traces the development of atomic ideas and models, introduces electrons, protons and neutrons, explains atomic number, mass number and electronic configuration and develops the ideas of valency, isotopes, average atomic mass and isobars.

1. Rediscovering the Roots of Atomic Theory

- More than 2,000 years ago, thinkers in ancient India and ancient Greece asked what matter is ultimately made of.
- Acharya Kanada proposed that repeated division of matter (dravya) would finally lead to indivisible particles called parmanus.
- His ideas are recorded in the Vaisesika Sutras. Combinations of parmanus were described as forming larger units.
- Greek philosophers Leucippus and Democritus proposed a similar idea and used the word atomos, meaning indivisible.
- These early ideas were philosophical rather than conclusions drawn from modern scientific experiments.
- In 1808, John Dalton proposed an experimental scientific atomic theory in which matter was composed of indivisible atoms.

2. Why Atomic Models Changed

- Scientific models of the atom changed whenever new experiments produced evidence that older models could not explain.
- Radioactivity and the discovery of subatomic particles showed that atoms were not indivisible.
- The development of atomic models illustrates how scientific understanding advances through observation, questioning and experimentation.

3. Discovery of the Electron

- In 1897, J. J. Thomson studied the conduction of electricity through gases at very low pressure using a cathode ray tube.
- Cathode rays travelled from the cathode (negative electrode) toward the anode (positive electrode).
- Their behaviour in electric and magnetic fields showed that they were streams of negatively charged particles.
- These particles, later called electrons, had much smaller mass than atoms.
- Cathode rays were independent of the cathode material and gas used, showing that electrons are present in all atoms.
- The relative charge of an electron is taken as -1.

4. Thomson's Model of the Atom

Plum pudding model
Thomson proposed that an atom is a sphere of positive charge with negatively charged electrons distributed throughout it.

- The model was compared with plums embedded in a pudding or seeds distributed through the pulp of a watermelon.
- It attempted to explain why an atom is electrically neutral: total positive and negative charges balance.
- Later experiments showed that the positive charge is not spread uniformly throughout the atom.

5. Rutherford's Gold Foil Experiment

- In 1911, Geiger and Marsden, working under Ernest Rutherford, directed a narrow beam of positively charged alpha (α) particles at a very thin gold foil.
- According to Thomson's model, the particles were expected to pass through or suffer only small deflections.

Observation	**What it suggested**
Most α-particles passed straight through	Most of the atom is empty space.
Some α-particles were strongly deflected	Positive charge is concentrated in a small region.
A very few bounced back	The central region is extremely small, dense and massive.

- The deflection of particles from a straight path is called scattering; therefore, the experiment is also called the α-ray scattering experiment.
- Thomson's model could not explain the large-angle deflections and back-scattering.

6. Rutherford's Model of the Atom

- The positive charge is concentrated in a tiny central region called the nucleus.
- The nucleus is dense and contains all the positive charge and most of the mass of the atom.
- Most of the atom is empty space.
- Electrons revolve around the nucleus, somewhat like planets around the Sun; hence it is called the planetary model.
- The diameter of an atom is about 10^{-10} m, while the diameter of its nucleus is about 10^{-15} m.

7. Limitation of Rutherford's Model

- Rutherford's model could not explain the stability of atoms.
- An electron moving in a circular path is accelerating because its direction continuously changes.
- According to the reasoning, an accelerating charged electron should lose energy, spiral inward and finally fall into the positively charged nucleus.
- Real atoms are stable, so a new model was needed.

8. Discovery of the Proton

- Rutherford showed that the positive charge of the nucleus comes from particles called protons.
- A proton is much heavier than an electron and has a relative charge of +1.
- For an electrically neutral atom, number of protons = number of electrons.
- Thus, the total positive charge and total negative charge cancel each other.

9. Bohr's Model of the Atom

- Niels Bohr proposed a new atomic model in 1913 to explain atomic stability.
- Electrons move only in fixed circular paths called stationary states orbits or shells.
- Each shell has a definite energy and is therefore also called an energy level.
- Shells are represented as K, L, M, N, ... or by n = 1, 2, 3, 4, ...
- K-shell (n = 1) is nearest the nucleus and has the least energy.
- Energy increases as the distance of the shell from the nucleus increases.
- An electron in an allowed stationary shell does not lose energy.
- An electron can move between shells only by absorbing or releasing a fixed amount of energy equal to the energy difference between the levels.
- Each shell can hold only a limited number of electrons.

Further development

The Bohr's model also has limitations. Modern quantum theory describes electrons in terms of regions of probability or 'electron clouds', rather than fixed classical paths.

10. Subatomic Particles

Particle	Symbol	Relative charge	Main location / role
Electron	e^-	-1	Outside nucleus; extremely small mass
Proton	p^+	+1	Nucleus; determines atomic number
Neutron	n^0	0	Nucleus; contributes substantially to atomic mass

11. Discovery of the Neutron

- James Chadwick discovered the neutron in 1932.
- A neutron has no electrical charge and has a mass nearly equal to that of a proton.
- Neutrons occur in the nuclei of atoms except ordinary hydrogen.
- Most of the mass of an atom comes from protons and neutrons packed in the nucleus.
- Protons and neutrons together are called nucleons.

12. Symbols of Elements

- Dalton introduced pictorial symbols for elements, but later alphabetic symbols became standard.
- In 1813, Berzelius suggested symbols derived from element names, including Latin names.
- Today, IUPAC approves the names and symbols of elements.
- The first letter of a chemical symbol is always capital; a second letter, if present, is lowercase.
- Examples: H, Al, Co, Cl and Zn.
- Some symbols come from Latin, Greek or German names: Fe from ferrum, Hg from hydrargyros, W from wolfram, Na from natrium, K from kalium and Ag from argentum.

Element	Symbol	Element	Symbol
Hydrogen	H	Oxygen	O
Carbon	C	Nitrogen	N
Sodium	Na	Potassium	K
Iron	Fe	Copper	Cu
Silver	Ag	Gold	Au
Lead	Pb	Mercury	Hg
Chlorine	Cl	Magnesium	Mg
Aluminium	Al	Silicon	Si

13. Atomic Number

Definition
The number of protons in the nucleus of an atom is called its atomic number. It is represented by Z.

- Atomic number determines the identity of an element.
- For a neutral atom: number of protons = number of electrons = atomic number.
- Example: hydrogen has Z = 1; helium has Z = 2.

14. Mass Number

Definition
The total number of protons and neutrons in the nucleus is called the mass number. It is represented by A.

- Mass number A = number of protons + number of neutrons.
- Number of neutrons = A - Z.
- The mass of electrons is almost negligible in these calculations.
- Standard atomic notation places mass number A at the upper left and atomic number Z at the lower left of the element symbol.

Element	Protons	Neutrons	Mass number
Hydrogen	1	0	1
Helium	2	2	4
Lithium	3	4	7
Carbon	6	6	12

15. Distribution of Electrons in Shells

- Bohr and Bury gave rules for distributing electrons in energy levels.
- Maximum electrons in a shell = $2n^2$, where n is the shell number.
- Thus K (n=1) can hold 2, L (n=2) can hold 8 and M (n=3) can theoretically hold 18 electrons.
- The outermost shell can accommodate a maximum of 8 electrons; the first shell can accommodate a maximum of 2.
- Electrons fill shells stepwise from inner to outer: K, then L, then M, then N, …
- The arrangement of electrons among shells is called electronic configuration.

16. Electronic Configuration of the First 18 Elements

Element	Z	Electronic configuration
H	1	1
He	2	2
Li	3	2, 1
Be	4	2, 2
B	5	2, 3
C	6	2, 4
N	7	2, 5
O	8	2, 6
F	9	2, 7
Ne	10	2, 8
Na	11	2, 8, 1
Mg	12	2, 8, 2
Al	13	2, 8, 3
Si	14	2, 8, 4
P	15	2, 8, 5
S	16	2, 8, 6
Cl	17	2, 8, 7
Ar	18	2, 8, 8

17. Valence Shell and Valence Electrons

- The outermost occupied shell of an atom is called its valence shell.
- Electrons present in the valence shell are called valence electrons.
- An outermost shell containing 8 electrons is called an octet.
- Atoms with a complete octet or helium with two electrons in its only shell, are largely stable and unreactive.
- Atoms with incomplete valence shells generally tend to lose, gain or share electrons to reach a stable configuration.

18. Valency

Definition
Valency is the combining capacity of an atom. It is the number of electrons gained, lost or shared by an atom to achieve a stable electronic configuration.

- If an atom has fewer than four valence electrons, it generally tends to lose electrons to complete its octet.
- If it has more than four valence electrons, it generally tends to gain electrons.
- An atom with four valence electrons can complete its octet by sharing electrons.

Element	Configuration	Valence electrons	Common valency
Sodium	2, 8, 1	1	1
Magnesium	2, 8, 2	2	2
Carbon	2, 4	4	4
Oxygen	2, 6	6	2
Chlorine	2, 8, 7	7	1
Neon	2, 8	8	0

19. Isotopes

Definition
Isotopes are atoms of the same element having the same atomic number but different mass numbers.

- Isotopes have the same number of protons but different numbers of neutrons.
- Because they have the same number of electrons and the same electronic configuration, their chemical properties are similar.
- Their physical properties can differ.
- Hydrogen has three isotopes: protium (^1H), deuterium (^2H) and tritium (^3H).
- Carbon has isotopes ^{12}C, ^{13}C and ^{14}C.

20. Uses of Isotopes

Isotope	Uses
Uranium-235	Fuel in nuclear reactors for electricity generation
Cobalt-60	Radiation treatment for cancer
Iodine-131	Treatment of goitre and thyroid cancer
Carbon-14	Determining the age of ancient fossils and artefacts

21. Average Atomic Mass

- Natural samples of an element may contain isotopes in different proportions.
- Therefore, average atomic mass must take the relative abundance of each isotope into account.
- This gives a weighted average atomic mass rather than a simple arithmetic mean.

Example: chlorine
Her, the chlorine-35 (about 75%) and chlorine-37 (about 25%).
Weighted average atomic mass = (35 × 75/100) + (37 × 25/100) = 35.5 u.

- A value such as 35.5 u does not mean that one chlorine atom has a fractional mass of 35.5 u; it is an average for a natural sample.

22. Isobars

Definition
Isobars are atoms of different elements having the same mass number but different atomic numbers.

- Example: argon, potassium and calcium can have mass number 40 but different atomic numbers.
- Isotopes: same Z, different A. Isobars: same A, different Z.

23. Development of Atomic Models - Timeline

Model / idea	Main contribution
Kanada / Democritus	Early philosophical idea of indivisible particles
Dalton	Atoms described scientifically as indivisible building blocks of matter
Thomson	Electrons embedded in a positively charged sphere
Rutherford	Tiny dense positive nucleus; atom mostly empty space
Bohr	Electrons in fixed energy levels or shells
Modern model	Electron probability clouds; quantum mechanical description

24. Important Relations

Quantity	Relation
Atomic number	Z = number of protons
Neutral atom	protons = electrons = Z
Mass number	A = protons + neutrons
Neutrons	n = A - Z
Maximum electrons in nth shell	$2n^2$
Isotopes	Same Z, different A
Isobars	Same A, different Z

25. Common Conceptual Points

- An atom is not indivisible; it contains subatomic particles.
- Most of the atom is empty space, while most of its mass is concentrated in the nucleus.
- Atomic number, not mass number, uniquely identifies an element.
- In a neutral atom, the number of electrons equals the number of protons.
- Protons and neutrons are nucleons and contribute nearly all of the atom's mass.
- Electronic configuration determines the number of valence electrons and helps explain combining capacity.
- Atoms with complete outer shells are generally more stable and less reactive.
- Isotopes belong to the same element because they have the same atomic number.
- Isobars are different elements even though their mass numbers are equal.
- Scientific atomic models are models based on evidence and have been modified as new evidence emerged.

26. Quick Solved Examples

Example 1: Protons, electrons and neutrons

For an atom with Z = 12 and A = 24: protons = 12, electrons = 12 (neutral atom), neutrons = 24 - 12 = 12.

Example 2: Mass number

An atom has 17 protons and 18 neutrons. A = 17 + 18 = 35.

Example 3: Electronic configuration

Atomic number 16 means 16 electrons in a neutral atom. Electronic configuration = 2, 8, 6. It has 6 valence electrons and common valency 2.

Example 4: Isotope check

Two atoms each have 11 protons, but one has 12 neutrons and the other 13. Both have Z = 11, while A = 23 and 24. Therefore, they are isotopes of the same element.

27. Key Terms

Term	Meaning
Atom	Basic building block of matter.
Electron	Negatively charged subatomic particle.
Proton	Positively charged subatomic particle in the nucleus.
Neutron	Neutral subatomic particle in the nucleus.
Nucleus	Small dense central region containing protons and neutrons.
Shell / energy level	Allowed energy region for electrons in Bohr's model.
Atomic number	Number of protons in an atom's nucleus.
Mass number	Total number of protons and neutrons.
Nucleons	Protons and neutrons together.
Electronic configuration	Distribution of electrons among shells.
Valence shell	Outermost occupied shell.
Valence electrons	Electrons in the valence shell.
Valency	Combining capacity of an atom.
Isotope	Same element: same Z, different A.
Isobar	Different elements: same A, different Z.

28. Exam-Oriented Self-Check Questions

1. How did the ideas of Acharya Kanada and Democritus contribute to early atomic thought?
2. Why was Dalton's atomic theory important in the development of atomic science?
3. What observations led Thomson to conclude that electrons are present in all atoms?
4. Explain Thomson's model of the atom.
5. State the main observations of Rutherford's gold foil experiment and the conclusions drawn from them.
6. Why did Rutherford conclude that most of an atom is empty space?
7. State the main features and limitation of Rutherford's atomic model.
8. How did Bohr's model explain the stability of an atom?
9. Compare electron, proton and neutron on the basis of charge and location.
10. Define atomic number and mass number.
11. An atom has atomic number 26 and mass number 56. Find its protons, electrons and neutrons.
12. State the Bohr-Bury rules for distribution of electrons in shells.
13. Write the electronic configurations of sodium, magnesium, oxygen, chlorine and argon.
14. Define valence shell, valence electrons and valency.
15. Why are noble-gas-like complete outer shells considered stable?
16. Define isotopes with examples.
17. Why do isotopes of an element have similar chemical properties?
18. Give four uses of isotopes.
19. What is weighted average atomic mass? Explain using chlorine as an example.
20. Differentiate between isotopes and isobars with examples.

One-Minute Chapter Summary

Quick revision

Atoms are the building blocks of matter, but they contain smaller particles. Thomson discovered the electron and proposed a positively charged sphere containing electrons. Rutherford's gold foil experiment revealed that an atom is mostly empty space with a tiny, dense, positively charged nucleus. Bohr proposed fixed energy levels or shells. Protons and neutrons lie in the nucleus, while electrons occupy shells. Atomic number Z equals the number of protons; mass number A equals protons plus neutrons. Electrons are arranged according to shell-filling rules and valence electrons determine valency. Isotopes have the same atomic number but different mass numbers, while isobars have the same mass number but different atomic numbers. Natural isotopic abundance is used to calculate weighted average atomic mass.

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