



STUDY MATERIAL

JOURNEY INSIDE THE ATOM

Grade 9 Science | Theory Notes • Solved Examples • Practice Questions

IN THIS BOOKLET

Concise theory notes • Important definitions and formulas • Worked examples • Practice set • Answer key

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Prepared from the chapter PDF supplied. These notes reorganise and simplify the chapter for revision; use them alongside your textbook and teacher's explanations.

1. Roots of Atomic Theory

Acharya Kanada: proposed that repeated division of matter would eventually reach tiny indivisible particles called *parmanus*. His ideas are recorded in the *Vaisesika Sutras*.

Leucippus and Democritus: suggested indivisible particles called *atomos*, a Greek word meaning indivisible.

John Dalton (1808): proposed a scientific atomic theory based on experimental evidence available at the time. He considered atoms indivisible; later discoveries established that atoms contain smaller particles.

Scientific models

Models change when new experiments provide evidence that an existing explanation cannot account for.

2. Atomic Models

Thomson's model

- J. J. Thomson studied cathode rays in a low-pressure tube and concluded that they are streams of negatively charged particles, later called electrons.
- He pictured an atom as a sphere of positive charge with electrons embedded throughout it—the plum-pudding model. A watermelon analogy uses pulp for positive charge and seeds for electrons.

Rutherford's gold foil experiment (1911)

- Geiger and Marsden, working under Rutherford, directed positively charged alpha particles at thin gold foil.
- **Observation:** most passed straight through. **Conclusion:** most of an atom is empty space.
- **Observation:** some were deflected. **Conclusion:** positive charge is concentrated in a small region.
- **Observation:** a very few bounced back. **Conclusion:** the nucleus is tiny, dense and contains most of the atom's mass.

Rutherford's model placed electrons around a central nucleus. Its limitation was that it could not explain atomic stability: classical reasoning suggested orbiting electrons would lose energy and spiral into the nucleus.

Bohr's model (1913)

- Electrons occupy allowed shells or energy levels: K, L, M, N... or $n = 1, 2, 3, 4...$
- Electrons do not lose energy while remaining in an allowed stationary shell.
- The K shell is closest to the nucleus and has the least energy. Shell energy increases farther from the nucleus.
- Electrons absorb or release a fixed amount of energy when they move between levels.

Modern model: the chapter introduces the quantum-mechanical model, where electrons are described as electron clouds—regions where they are most likely to be found—rather than exact fixed paths.

3. Compare the Atomic Models

Model	Main idea	Limitation / next development
Dalton	Atoms as indivisible particles	Subatomic particles were discovered later.
Thomson	Electrons embedded in a positive sphere	Could not explain alpha-particle scattering.
Rutherford	Tiny positive nucleus; atom mostly empty space	Could not explain atomic stability.
Bohr	Electrons in fixed energy shells	Later developed into the quantum-mechanical model.

Gold foil experiment: quick memory link

Most pass through → mostly empty space. A few deflect strongly → concentrated positive nucleus. A very few bounce back → nucleus is tiny and dense.

4. Subatomic Particles

Particle	Symbol	Relative charge	Location
Electron	e ⁻	-1	Outside the nucleus
Proton	p ⁺	+1	Nucleus
Neutron	n ⁰	0	Nucleus

- A neutral atom has equal numbers of protons and electrons.
- James Chadwick discovered the neutron in 1932. Neutrons have nearly the same mass as protons but no electrical charge.
- The nucleus contains protons and neutrons; electrons have very small mass and are usually ignored in mass-number calculations.

5. Symbols of Elements

- The first letter of a chemical symbol is uppercase; a second letter, if present, is lowercase: H, C, Al, Co.
- Some symbols come from Latin or other names: iron Fe (ferrum), sodium Na (natrium), potassium K (kalium), copper Cu (cuprum), silver Ag (argentum), gold Au (aurum), lead Pb (plumbum), tungsten W (wolfram).

Common symbols to learn: Hydrogen H; Oxygen O; Nitrogen N; Carbon C; Sodium Na; Potassium K; Calcium Ca; Magnesium Mg; Aluminium Al; Silicon Si; Chlorine Cl; Sulfur S; Iron Fe; Copper Cu; Silver Ag; Gold Au; Lead Pb; Zinc Zn.

6. Atomic Number and Mass Number

Atomic number (Z) is the number of protons in the nucleus. It identifies the element. For a neutral atom, electrons = protons = Z.

Mass number (A) is the total number of protons and neutrons in the nucleus. Protons and neutrons are called nucleons.

Formula box

Z = number of protons

Neutral atom: electrons = Z

A = protons + neutrons

Neutrons = A – Z

Standard notation

An atom is written as ${}^A_Z\text{X}$: A (mass number) is at the upper left; Z (atomic number) is at the lower left. Example: carbon-12 = ${}^{12}_6\text{C}$.

Solved examples

Example 1: Z = 17 and A = 35. Protons = 17; electrons in a neutral atom = 17; neutrons = 35 – 17 = **18**.

Example 2: An atom has 12 protons and 13 neutrons. Z = **12**; A = 12 + 13 = **25**.

Example 3: An atom has 79 protons and A = 197. Neutrons = 197 – 79 = **118**; neutral atom electrons = **79**.

7. Electronic Configuration

The arrangement of electrons among shells is called electronic configuration. Shells fill stepwise, starting with the shell closest to the nucleus.

- Maximum shell capacity = $2n^2$, where n is the shell number.
- K shell (n=1): 2 electrons; L shell (n=2): 8; M shell (n=3): maximum capacity 18 by the formula.
- For the first 18 elements, the configurations are built using K, L and M shells; the outermost shell holds no more than 8 electrons (the first shell holds at most 2).

Electronic Configuration: First 18 Elements

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

Example

Magnesium has $Z = 12$, so a neutral atom has 12 electrons. Configuration = 2, 8, 2.

8. Valency

Valence shell: outermost occupied shell. **Valence electrons:** electrons in that shell. **Valency** is an atom's combining capacity; in this chapter it is the number of electrons gained, lost or shared to achieve a stable outer shell.

- A full outer shell is generally stable: 8 electrons (octet), or 2 for helium.
- With 1–3 valence electrons, atoms commonly lose electrons in the simple model; with 5–7, they commonly gain or share electrons. Atoms with 4 valence electrons may share four.

Examples

Sodium: 2, 8, 1 → loses one electron → valency 1. Oxygen: 2, 6 → needs two electrons to complete an octet → valency 2. Carbon: 2, 4 → shares four → valency 4. Neon: 2, 8 → stable shell → valency 0 in this introductory model.

9. Isotopes, Isobars and Average Atomic Mass

Isotopes: atoms of the same element with the same atomic number but different mass numbers. They have the same number of protons but different numbers of neutrons. Their chemical properties are similar because their electronic configurations are the same.

Hydrogen isotopes: protium ${}^1\text{H}$ (0 neutrons), deuterium ${}^2\text{H}$ (1 neutron), tritium ${}^3\text{H}$ (2 neutrons).

Isobars: atoms of different elements with the same mass number but different atomic numbers. Example: calcium-40 ($Z=20$), potassium-40 ($Z=19$), argon-40 ($Z=18$).

Uses of isotopes listed in the chapter

- Uranium-235: fuel in nuclear reactors.
- Cobalt-60: radiation treatment for cancer.
- Iodine-131: treatment of goitre and thyroid cancer.
- Carbon-14: estimating the age of ancient fossils and artefacts.

Weighted average atomic mass

Average atomic mass accounts for each isotope's mass and its natural abundance. Formula: average mass = $\Sigma(\text{isotope mass} \times \text{fractional abundance})$.

Formula box

Average atomic mass = $(m_1 \times \% \text{abundance}_1/100) + (m_2 \times \% \text{abundance}_2/100) + \dots$

Worked example: chlorine — mass 35 u at 75%, mass 37 u at 25%. Average = $(35 \times 0.75) + (37 \times 0.25) = 26.25 + 9.25 = 35.5 \text{ u}$.

Worked example: bromine — 79 u at 49.7%, 81 u at 50.3%. Average = $(79 \times 0.497) + (81 \times 0.503) = 80.006 \text{ u} \approx 80.0 \text{ u}$.

10. Practice Questions

A. Multiple choice

1. Who proposed the plum-pudding model? (A) Dalton (B) Thomson (C) Rutherford (D) Bohr
2. Most alpha particles passing through gold foil indicates that an atom is mostly: (A) solid (B) empty space (C) neutrons (D) negative charge
3. Which particle has no electrical charge? (A) Proton (B) Electron (C) Neutron (D) Alpha particle
4. Atomic number equals the number of: (A) neutrons (B) protons (C) nucleons (D) shells
5. Atoms with the same Z but different A are: (A) isobars (B) isotopes (C) ions (D) molecules
6. Correct configuration for sodium ($Z=11$): (A) 2,9 (B) 2,8,1 (C) 2,7,2 (D) 8,3
7. Correct symbol for cobalt: (A) CO (B) co (C) Co (D) Cb

Practice Questions (continued)

B. Short-answer questions

8. Define atomic number and mass number.
9. State three observations from Rutherford's gold foil experiment and the conclusion from each.
10. Give one limitation each of Thomson's and Rutherford's models.
11. State Bohr's main ideas about electron shells.
12. Why do isotopes of the same element have similar chemical properties?
13. Distinguish between isotopes and isobars with one example of each.
14. Why is weighted average atomic mass more accurate than a simple average?
15. Arrange Dalton, Thomson, Rutherford and Bohr models in chronological order.

C. Numerical and application questions

16. A neutral atom has $Z=13$ and $A=27$. Find protons, electrons and neutrons.
17. An atom has 20 protons and 22 neutrons. Find Z and A .
18. An atom has $A=40$ and 18 electrons. Assuming it is neutral, find protons and neutrons.
19. Write configurations for atomic numbers 8, 12, 15 and 18.
20. Find valency for configurations (a) 2,8,1 (b) 2,6 (c) 2,8,7 (d) 2,8,8.
21. Two atoms each have 11 protons; one has 12 neutrons and the other 13. Find their mass numbers and relationship.
22. An element has isotopes of masses 10 u (20%) and 11 u (80%). Calculate average atomic mass.
23. Bromine has isotopes 79 u (49.7%) and 81 u (50.3%). Calculate average atomic mass.
24. An atom has 17 protons and 18 neutrons. Find Z , A and electrons if neutral.
25. An atom is represented as $^{23}_{11}\text{X}$. Identify X and calculate protons, electrons and neutrons.
26. Elements have (Z,A) values (18,40) and (20,40). Are they isotopes or isobars? Explain.
27. A neutral atom has 10 electrons and 12 neutrons. Identify it and find its mass number.

Exam technique

Write the formula before substituting values. For atomic structure, always distinguish atomic number (protons) from mass number (protons + neutrons). For experiments, connect each observation to its conclusion.

11. Answer Key

A. Multiple choice

1. B • 2. B • 3. C • 4. B • 5. B • 6. B • 7. C

B. Short answers — key points

8. Z = number of protons. A = protons + neutrons.

9. Most passed through → mostly empty space; some deflected → positive charge concentrated in a small region; a few bounced back → tiny, dense nucleus containing most of the mass.

10. Thomson's model could not explain large alpha-particle deflections. Rutherford's model could not explain atomic stability.

11. Electrons occupy allowed shells of fixed energy; they do not lose energy while in a stationary shell; they absorb/release a fixed amount when changing shells.

12. Same number of electrons and same electronic configuration; chemical properties depend mainly on electrons, especially valence electrons.

13. Isotopes: same Z , different A (e.g. ^{12}C and ^{13}C). Isobars: same A , different Z (e.g. ^{40}Ar and ^{40}Ca).

14. Weighted average accounts for natural abundance; simple average assumes all isotopes are equally abundant.

15. Dalton → Thomson → Rutherford → Bohr.

C. Numerical answers

16. Protons 13; electrons 13; neutrons $27 - 13 = 14$.

17. $Z=20$; $A=20+22 = 42$.

18. Protons 18; neutrons $40 - 18 = 22$.

19. $Z=8$: 2,6; $Z=12$: 2,8,2; $Z=15$: 2,8,5; $Z=18$: 2,8,8.

20. (a) 1 (b) 2 (c) 1 (d) 0.

21. Mass numbers 23 and 24; they are isotopes.

22. $(10 \times 0.20) + (11 \times 0.80) = 10.8 \text{ u}$.

23. $(79 \times 0.497) + (81 \times 0.503) = 80.006 \text{ u} \approx 80.0 \text{ u}$.

24. $Z=17$; $A=35$; neutral electrons=17.

25. X = sodium (Na); protons 11; electrons 11; neutrons $23 - 11 = 12$.

26. Isobars: same mass number 40 but different atomic numbers.

27. $Z=10$, element neon (Ne); $A=10+12=22$.

Final revision checklist

- Compare Dalton, Thomson, Rutherford and Bohr models.
- Explain the gold foil experiment.
- Calculate protons, electrons, neutrons, Z and A .
- Write configurations for the first 18 elements.
- Find common valencies; distinguish isotopes and isobars.
- Calculate weighted average atomic mass.

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