



CLASS 9 SCIENCE — CHAPTER 8

JOURNEY INSIDE THE ATOM



New Exploration Science Textbook | 2026–27

Chapter 8 takes us inside the atom and explains how scientists discovered **electrons, protons and neutrons**, and how atomic models developed from **Dalton → Thomson → Rutherford → Bohr**. It also covers **electronic configuration, atomic number, mass number, valency, isotopes and isobars**. ([Learn CBSE](#))



Formula format: All formulas are written in simple **text format**.



1. What Is an Atom?

Everything around us is made up of **matter**.

Matter is made up of extremely small particles called:



Atoms

An atom is the basic unit of an element that retains the chemical identity of that element.

Examples:



Hydrogen → Hydrogen atoms



Oxygen → Oxygen atoms



Carbon → Carbon atoms



2. Can We See an Atom?

Atoms are extremely small.

They cannot normally be seen with our naked eyes or an ordinary optical microscope.

Scientists learned about atoms through:



Experiments



Observations



Reasoning



Electrical experiments



Experiments involving radiation

Our understanding of the atom changed as new experimental evidence became available.



3. Development of Atomic Models

The modern atomic model was not developed in one step.

Scientists proposed different models based on experimental evidence.

Important sequence:

Dalton → Thomson → Rutherford → Bohr → Modern understanding

Let's understand each model. 📌



4. Dalton's Atomic Theory



John Dalton

Dalton proposed an early scientific atomic theory.

According to Dalton:

- 1 Matter is made up of tiny particles called atoms.**
- 2 Atoms of the same element were considered identical.**
- 3 Atoms of different elements have different properties.**
- 4 Atoms combine in simple whole-number ratios to form compounds.**

5 Chemical reactions involve rearrangement of atoms.

5. Dalton's Model

Dalton imagined an atom as:

A tiny, solid, indivisible sphere.



There were no known smaller particles inside it at that time.

★ Limitation

Later experiments showed that atoms are **not indivisible**.

They contain smaller particles called:

-  Electrons
-  Protons
-  Neutrons

Therefore, Dalton's model was replaced by newer models.

6. Discovery of the Electron

 **J. J. Thomson**

Thomson carried out experiments using:

Cathode rays

He found that cathode rays consisted of negatively charged particles.

These particles were later called:

 **Electrons**

7. Properties of Electrons

An electron:

- ♦ Has a **negative charge**
- ♦ Has a very small mass
- ♦ Is present outside the nucleus in an atom

Symbol:

e^-

8. Thomson's Atomic Model

Thomson proposed the:

Plum Pudding Model

According to this model:

● The atom was considered a positively charged sphere.

⚡ Electrons were embedded inside it.

The positive charge balanced the negative charge of the electrons.

Therefore, the atom as a whole was electrically neutral.

9. Why Was It Called the Plum Pudding Model?

Imagine:

 A positively charged pudding

with

 Small negative electrons embedded in it.

Hence:

Plum Pudding Model

10. Limitation of Thomson's Model

Thomson's model could not explain later experimental observations about the internal structure of the atom.

This led to another important experiment.

11. Rutherford's Experiment

 **Ernest Rutherford**

Rutherford and his collaborators performed the famous:

Alpha-particle scattering experiment

They directed alpha particles towards a very thin sheet of gold foil.

12. Why Gold Foil?

Gold can be beaten into extremely thin sheets.

This allowed alpha particles to pass through a very thin layer of atoms.

13. Rutherford's Observations

Most alpha particles:

➡ **Passed straight through the foil.**

Some:

↗ **Were deflected slightly.**

A very small number:

↶ **Were deflected through large angles.**

A tiny fraction:

← **Returned almost in the direction from which they came.**

These observations were extremely important.



14. Rutherford's Conclusions

From the experiment, Rutherford concluded:

① **Most of the atom is empty space.**

Because most alpha particles passed straight through.

② **Positive charge is concentrated in a very small region.**

③ **Most of the mass of the atom is concentrated in this small region.**

This central region was called:

Nucleus

④ **Electrons are present outside the nucleus.**



15. Rutherford's Atomic Model

According to Rutherford:

Nucleus

Contains most of the atom's mass and positive charge.

Electrons

Move around the nucleus.

Empty space

Most of the atom consists of empty space.

16. Limitation of Rutherford's Model

Rutherford's model could not explain why electrons do not continuously lose energy and spiral into the nucleus.

This problem was addressed by:

Niels Bohr

17. Bohr's Atomic Model

Bohr proposed that electrons occupy specific:

Energy levels or shells

around the nucleus.

Electrons cannot have arbitrary energies.

They occupy allowed energy levels.

18. Electron Shells

The electron shells are commonly represented as:

K, L, M, N

From the nucleus outward:

K → L → M → N

These correspond to increasing principal energy levels.

19. Electron Distribution

For the first few shells, the maximum number of electrons that can occupy a shell is given by:

Maximum electrons = $2n^2$

Where:

n = shell number

20. Maximum Electrons in K Shell

K shell:

n = 1

Using:

$2n^2$

= 2×1^2

= 2

Maximum = 2 electrons



21. Maximum Electrons in L Shell

L shell:

$$n = 2$$

$$2n^2 = 2 \times 2^2$$

$$= 2 \times 4$$

$$= 8$$

Maximum = 8 electrons



22. Maximum Electrons in M Shell

M shell:

$$n = 3$$

$$2n^2 = 2 \times 3^2$$

$$= 2 \times 9$$

$$= 18$$

Maximum = 18 electrons



23. Maximum Electrons in N Shell

N shell:

$$n = 4$$

$$2n^2 = 2 \times 4^2$$

$$= 2 \times 16$$

$$= 32$$

Maximum = 32 electrons



24. Shell Capacity

Shell	n	Maximum electrons
-------	---	-------------------

K	1	2
---	---	---

L	2	8
---	---	---

M	3	18
---	---	----

N	4	32
---	---	----



25. Structure of an Atom

An atom contains three important subatomic particles:

 **Electron**

Charge = Negative

 **Proton**

Charge = Positive

 **Neutron**

Charge = No charge



26. Subatomic Particles

Particle	Symbol	Charge	Location
----------	--------	--------	----------

Electron	e^-	Negative	Outside nucleus
Proton	p^+	Positive	Nucleus
Neutron	n^0	Neutral	Nucleus

27. The Nucleus

The nucleus is the tiny central region of an atom.

It contains:

 **Protons**

and

 **Neutrons**

Together they are called:

Nucleons

28. Electrons

Electrons are negatively charged particles.

They occupy regions around the nucleus associated with different energy levels.

Important:

The number of electrons in a **neutral atom** is equal to the number of protons.

Therefore:

Number of electrons = Number of protons

for a neutral atom.

29. Protons

Protons are positively charged particles.

They are present inside the nucleus.

The number of protons determines the identity of an element.

This number is called:

Atomic number

30. Neutrons

Neutrons have:

No net electric charge

They are present in the nucleus along with protons.

Neutrons contribute significantly to the mass of an atom.

31. Atomic Number

Atomic number is represented by:

Z

Definition:

Atomic number is the number of protons present in the nucleus of an atom.

Formula:

Atomic Number = Number of Protons

or

$$Z = p$$

For a neutral atom:

$$Z = p = e$$



32. Example — Atomic Number

Suppose an atom has:

Protons = 11

Then:

Atomic Number = 11

Therefore, the element is:

Sodium (Na)



33. Mass Number

Mass number is represented by:

A

It is the total number of:

+ Protons

and

● Neutrons

Formula:

Mass Number = Number of Protons + Number of Neutrons

or:

$$A = p + n$$



34. Finding Number of Neutrons

From:

$$A = p + n$$

Therefore:

$$\text{Number of Neutrons} = A - Z$$

because:

$$p = Z$$

So:

$$n = A - Z$$



35. Example — Finding Neutrons

An atom has:

Atomic number = 17

Mass number = 35

Therefore:

$$\text{Number of neutrons} = 35 - 17$$

$$\text{Number of neutrons} = 18$$



36. Important Atomic Number Relations

For a neutral atom:

$$\text{Protons} = \text{Electrons}$$

And:

$$\text{Neutrons} = \text{Mass Number} - \text{Atomic Number}$$

Therefore:

$$A = p + n$$

$$Z = p$$

$$n = A - Z$$

37. Representation of an Atom

An atom is often represented as:

$${}^a_x\text{E}$$

where:

A = Mass number

Z = Atomic number

E = Symbol of element

For example:



Here:

$$A = 23$$

$$Z = 11$$



38. Reading $^{23}_{11}\text{Na}$

Mass number:

$$A = 23$$

Atomic number:

$$Z = 11$$

Protons:

$$p = 11$$

Electrons in neutral sodium:

$$e = 11$$

Neutrons:

$$n = A - Z$$

$$n = 23 - 11$$

$$n = 12$$



39. Electronic Configuration

The arrangement of electrons in different shells is called:

Electronic configuration

Electrons occupy shells according to their energies.

For many commonly studied elements, we can represent configurations using K, L, M and N shells.



40. Examples of Electronic Configuration

Hydrogen (H)

Atomic number = 1

Electrons = 1

Configuration:

1

Helium (He)

Atomic number = 2

Electrons = 2

Configuration:

2

Lithium (Li)

Atomic number = 3

Configuration:

2, 1

Carbon (C)

Atomic number = 6

Configuration:

2, 4

Oxygen (O)

Atomic number = 8

Configuration:

2, 6

Neon (Ne)

Atomic number = 10

Configuration:

2, 8

Sodium (Na)

Atomic number = 11

Configuration:

2, 8, 1

Chlorine (Cl)

Atomic number = 17

Configuration:

2, 8, 7



41. Valence Shell

The outermost occupied electron shell is called the:

Valence shell

The electrons present in this shell are:

Valence electrons

These electrons play an important role in chemical bonding and chemical reactions.

42. Valency

Definition

Valency is the combining capacity of an atom.

It is related to the number of electrons an atom loses, gains or shares to achieve a stable outer-shell arrangement.

43. Valency — Simple Rule

For many main-group elements:

If valence electrons are 1, 2, 3 or 4:

Valency is generally:

1, 2, 3 or 4

If valence electrons are 5, 6 or 7:

Valency is generally:

8 – number of valence electrons

If the outer shell is complete:

Valency = 0



44. Examples of Valency

Element	Configuration	Valence electrons	Valency
H	1	1	1
Li	2,1	1	1
Be	2,2	2	2
B	2,3	3	3
C	2,4	4	4
N	2,5	5	3
O	2,6	6	2
F	2,7	7	1
Ne	2,8	8	0
Na	2,8,1	1	1
Mg	2,8,2	2	2
Cl	2,8,7	7	1
Ar	2,8,8	8	0



45. Stable Electronic Configuration

Atoms tend to achieve a more stable outer electronic arrangement.

Noble gases have particularly stable outer shells.

Examples:

Helium

2 electrons

Neon

2, 8

Argon

2, 8, 8

These atoms generally have very low tendency to react compared with many other elements.



46. Ions

An atom can gain or lose electrons and become electrically charged.

Such a charged particle is called an:

Ion

+ 47. Cation

When an atom:

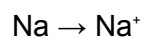
Loses electrons

it becomes positively charged.

This positive ion is called:

Cation

Example:



Sodium loses one electron.

— 48. Anion

When an atom:

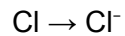
Gains electrons

it becomes negatively charged.

This negative ion is called:

Anion

Example:



Chlorine gains one electron.



49. Why Do Atoms Form Ions?

Atoms may gain or lose electrons to achieve a more stable electron arrangement.

For example:

Sodium

Configuration:

2, 8, 1

It can lose one electron:

2, 8

and form:

Na⁺

Chlorine

Configuration:

2, 8, 7

It can gain one electron:

2, 8, 8

and form:

Cl^-

50. Isotopes

Atoms of the same element have the same:

Atomic number

but may have different:

Mass numbers

Such atoms are called:

Isotopes

51. Example of Isotopes

Hydrogen has three commonly known isotopes:

^1H — Protium

^2H — Deuterium

^3H — Tritium

All have:

1 proton

but different numbers of neutrons.



52. Hydrogen Isotopes

Isotope	Protons	Neutrons	Mass number
^1H	1	0	1
^2H	1	1	2
^3H	1	2	3

★ **Remember:**

Same element → Same atomic number

Different isotopes → Different mass numbers



53. Isotopes of Carbon

Carbon has isotopes such as:

^{12}C

^{13}C

^{14}C

All have:

6 protons

but different numbers of neutrons.



54. Uses of Isotopes

Some isotopes have important applications.

Medicine

Certain radioactive isotopes are used in diagnosis and treatment.

Scientific research

Isotopes can be used as tracers.

Agriculture

Some isotopes are used in research related to plants and soil.

Archaeology

Carbon-14 is used for radiocarbon dating of suitable once-living materials.

55. Isobars

Atoms of different elements can have the same:

Mass number

but different:

Atomic numbers

Such atoms are called:

Isobars

56. Example of Isobars

Consider:

$^{40}_{18}\text{Ar}$

and



Both have:

Mass number = 40

But:

Argon atomic number = 18

Calcium atomic number = 20

Therefore, they are:

Isobars



57. Isotopes vs Isobars

Isotopes	Isobars
Same element	Different elements
Same atomic number	Different atomic numbers
Different mass numbers	Same mass number
Example: ^{12}C , ^{13}C , ^{14}C	Example: ^{40}Ar , ^{40}Ca



Memory Trick

ISO-TOPES → Same atomic number

ISO-BARS → Same mass number



58. Atomic Number vs Mass Number

Atomic Number

Mass Number

Symbol: Z

Symbol: A

Number of protons

Protons + neutrons

Identifies the element

Gives total nucleons

$Z = p$

$A = p + n$



59. Complete Structure of an Atom

Centre:

 **Nucleus**

Contains:

 Protons

 Neutrons

Outside nucleus:

 Electrons

arranged in different energy levels.



60. Subatomic Particles — Quick Revision

Particle	Charge	Relative location
 Electron	-1	Outside nucleus
 Proton	+1	Nucleus
 Neutron	0	Nucleus



61. Important Formula Sheet

Atomic number

$Z = \text{Number of protons}$

For a neutral atom

$\text{Number of electrons} = \text{Number of protons}$

Therefore:

$$e = p = Z$$

Mass number

$A = \text{Number of protons} + \text{Number of neutrons}$

Number of neutrons

$$n = A - Z$$

Maximum electrons in shell

$\text{Maximum electrons} = 2n^2$

where:

$n = \text{shell number}$



62. Numerical Example 1

An atom has:

Atomic number = 15

Mass number = 31

Find:

Protons
Electrons
Neutrons

Step 1: Protons

$$p = Z = 15$$

Step 2: Electrons

Neutral atom:

$$e = p$$

$$e = 15$$

Step 3: Neutrons

$$n = A - Z$$

$$n = 31 - 15$$

$$n = 16$$

✓ **Answer:**

Protons = **15**

Electrons = **15**

Neutrons = **16**



63. Numerical Example 2

An atom has:

20 protons

18 electrons

20 neutrons

Is it neutral?

No ✖

Because:

Protons = 20

Electrons = 18

There are 2 more positive charges than negative charges.

Therefore:

Charge = +2

The particle is a:

2+ ion



64. Numerical Example 3

An atom has electronic configuration:

2, 8, 7

Find:

Atomic number

Valence electrons

Valency

Total electrons:

$2 + 8 + 7 = 17$

For a neutral atom:

Atomic number = 17

Valence electrons:

7

Valency:

$$8 - 7 = 1$$

 **Answer:**

Atomic number = 17

Valence electrons = 7

Valency = 1

The element is:

Chlorine



65. Numerical Example 4

An atom has:

Mass number = 39

Atomic number = 19

Find neutrons.

$$n = A - Z$$

$$n = 39 - 19$$

$$n = 20$$



66. Competency-Based Question

A neutral atom has 12 electrons.

Find its atomic number.

For a neutral atom:

Electrons = Protons

Therefore:

$$p = 12$$

And:

$$Z = p$$

Atomic number = 12

The element is:

Magnesium (Mg)

67. Competency-Based Question

An atom has 8 protons and 10 electrons.

Is it neutral?

 **No.**

Protons = 8

Electrons = 10

There are two extra electrons.

Therefore charge:

-2

It is an anion.

68. Competency-Based Question

Two atoms have:

Atom A:

$$Z = 17$$

$$A = 35$$

Atom B:

$$Z = 17$$

$$A = 37$$

Are they isotopes?

 **Yes.**

They have:

Same atomic number = 17

Different mass numbers = 35 and 37

Therefore:

They are isotopes of the same element.



69. Competency-Based Question

Two atoms have:

Atom A:

$$A = 40$$

$$Z = 18$$

Atom B:

$$A = 40$$

$$Z = 20$$

Are they isotopes?

✗ No.

They have the same mass number but different atomic numbers.

Therefore:

They are isobars.

70. Common Mistakes to Avoid

✗ Mistake 1

Atomic number = protons + neutrons.

✓ Correct

Atomic number = Number of protons

✗ Mistake 2

Mass number = protons only.

✓ Correct

Mass number = Protons + Neutrons

✗ Mistake 3

Isotopes have the same mass number.

✓ Correct

Isotopes have:

Same atomic number + Different mass numbers

Mistake 4

Isobars are atoms of the same element.

Correct

Isobars are atoms of:

Different elements + Same mass number

Mistake 5

Neutrons have a positive charge.

Correct

Neutrons are electrically neutral.

Mistake 6

Electrons are found inside the nucleus.

Correct

Protons and neutrons are in the nucleus; electrons occupy regions around the nucleus.



71. Important Exam Questions

Q1. What is an atom?

An atom is the basic unit of an element that retains the chemical identity of that element.

Q2. Name the three subatomic particles.

Electron, proton and neutron.

Q3. What is the charge of an electron?

Negative.

Q4. What is the charge of a proton?

Positive.

Q5. What is the charge of a neutron?

Zero.

Q6. What is atomic number?

The number of protons present in the nucleus.

Q7. What is mass number?

The total number of protons and neutrons in the nucleus.

Q8. What is valency?

The combining capacity of an atom.

Q9. What are isotopes?

Atoms of the same element having the same atomic number but different mass numbers.

Q10. What are isobars?

Atoms of different elements having the same mass number but different atomic numbers.

Q11. What did Rutherford's experiment show?

It showed that the atom has a tiny, dense, positively charged nucleus and that most of the atom is empty space.

Q12. What did Bohr propose?

He proposed that electrons occupy specific energy levels around the nucleus.

 72. QUICK REVISION

 Matter is made of atoms.

 Dalton → solid sphere model.

 Thomson → discovered electron.

 Thomson → plum pudding model.

 Rutherford → nucleus.

 Most of atom → empty space.

 Bohr → fixed energy levels/shells.

 Proton → positive.

 Neutron → neutral.

 Electron → negative.

 Nucleus → protons + neutrons.

 Atomic number → number of protons.

 Mass number → protons + neutrons.

 Neutrons = Mass number – Atomic number.

 Neutral atom → protons = electrons.

 Electronic configuration → arrangement of electrons in shells.

 Valence electrons → electrons in outermost shell.

 Valency → combining capacity.

 Cation → positive ion.

 Anion → negative ion.

 Isotopes → same Z, different A.

 Isobars → same A, different Z.



73. MEMORY TRICKS



ATOMIC NUMBER

Z = Protons

Think:

Z → Identity of the element



MASS NUMBER

A = Protons + Neutrons

Think:

A = All nucleons



NEUTRONS

n = A - Z



ATOM

Nucleus = Protons + Neutrons

Outside nucleus = Electrons



ISOTOPES

Same Z → Different A



ISOBARS

Same A → Different Z

🌟 74. CHAPTER 8 — FINAL TAKEAWAY

Our understanding of the atom developed through a series of scientific discoveries. 🔬

🌐 Dalton

Atom → Solid sphere



⚡ Thomson

Discovered electrons → Plum pudding model



🎯 Rutherford

Discovered the nuclear nature of the atom



🪐 Bohr

Electrons → Specific energy levels



🧬 Modern understanding

Atom contains:

Protons + Neutrons + Electrons

The most important relationships to remember are:

★ **Z = Number of protons**

★ **A = Protons + Neutrons**

★ **Neutrons = $A - Z$**

★ **For a neutral atom: Protons = Electrons**

★ **Maximum electrons in shell = $2n^2$**

And the most important comparison:

↺ **Isotopes** → **Same atomic number, different mass number**

⚖ **Isobars** → **Same mass number, different atomic number**

These concepts form the foundation for understanding the **structure of matter and chemical behaviour** in the chapters that follow.