



CLASS 9 SCIENCE — CHAPTER 9

ATOMIC FOUNDATIONS OF MATTER



New *Exploration* Science Textbook | 2026–27

This chapter connects **chemical combination** → **atoms** → **molecules** → **chemical bonding** → **ions** → **chemical formulae** → **molecular mass and formula unit mass**.



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



1. Introduction

We see many substances around us:



Salt



Water



Sugar



Minerals



Air



Medicines

When substances undergo chemical changes, they combine in definite ways.

To understand **why substances combine in particular ratios**, we need to understand:



Atoms



Molecules



Chemical bonds



Ions



Chemical formulae



2. Law of Conservation of Mass

Antoine Lavoisier

The **Law of Conservation of Mass** states:

Mass can neither be created nor destroyed in a chemical reaction.

Therefore:

Total mass of reactants = Total mass of products

This law was proposed by **Antoine Lavoisier**, who is often called the **Father of Modern Chemistry**.

3. Reactants and Products

The substances that take part in a chemical reaction are called:

Reactants

The new substances formed are called:

Products

For example:

Reactants → Products

The atoms present in the reactants are rearranged to form the products.

4. Example of Conservation of Mass

Suppose the masses of reactants are:

4.0 g + 2.92 g

Total reactant mass:

4.0 + 2.92 = 6.92 g

Suppose products have masses:

$$1.76 \text{ g} + 0.72 \text{ g} + 4.44 \text{ g}$$

Total product mass:

$$1.76 + 0.72 + 4.44 = 6.92 \text{ g}$$

Therefore:

Mass of reactants = Mass of products

$$6.92 \text{ g} = 6.92 \text{ g} \quad \checkmark$$

The chapter gives this type of numerical illustration of conservation of mass.



5. Closed System and Open System



Closed System

In a closed system:

Matter does not escape into the surroundings.

Therefore, the measured mass remains constant.



Open System

In an open system, a gas produced during a reaction may escape.

The measured mass may appear to decrease.



Important

The mass has **not been destroyed**.

Some matter has simply escaped from the system.

The chapter specifically compares open and closed arrangements to explain this apparent loss of mass.



6. Example — Vinegar and Baking Soda

When vinegar reacts with baking soda:



Vinegar + Baking soda



Carbon dioxide gas is produced.

If the reaction takes place in an:

Open container

CO₂ escapes into the surroundings.

The measured mass may decrease.

Closed container

The gas remains within the system.

Therefore, the total mass remains essentially constant.



7. Law of Constant Proportions

This is also called:

Law of Definite Proportions



Definition

A pure compound always contains the same elements combined in the same fixed ratio by mass, regardless of its source.

For example:

Water

Water always contains:

Hydrogen : Oxygen

in the mass ratio:

1 : 8

The source of water does not change this ratio.



8. Example — Water

Suppose:

Hydrogen = 20 g

The mass ratio is:

Hydrogen : Oxygen = 1 : 8

Therefore:

Oxygen required = 20×8

Oxygen = 160 g

Mass of water formed:

$20 + 160 = 180$ g



Water = 180 g



9. Example — Sodium Chloride

For sodium chloride:

Sodium : Chlorine = **23 : 35.5**

Suppose sodium = 46 g.

Chlorine required:

Chlorine = $(35.5 / 23) \times 46$

Chlorine = 71 g

Therefore:

Mass ratio remains fixed.

10. Compound vs Mixture

Compound

Elements are chemically combined in a:

Fixed ratio

Example:

 Water

Hydrogen and oxygen are always present in a definite mass ratio.

Mixture

Components can be present in:

Any proportion

Example:

Salt + water

Different amounts of salt can be added.

 **Remember:**

Compound → Fixed proportion

Mixture → Variable proportion



11. Dalton's Atomic Theory



John Dalton

Dalton proposed his atomic theory to explain the laws of chemical combination.

Important ideas include:

1

Matter is made of extremely small particles called:

Atoms

2

Atoms of the same element were considered to have the same properties.

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

Atoms are not created or destroyed during a chemical reaction.

12. Limitations of Dalton's Theory

Modern science has shown that some original statements need modification.

Atoms are not indivisible.

They contain:

-  Electrons
-  Protons
-  Neutrons

Also, atoms of the same element can have different numbers of neutrons.

These are called:

Isotopes

So Dalton's theory was an important foundation, but later discoveries refined our understanding.

13. Atom

 **Definition**

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

Examples:

H → Hydrogen atom

O → Oxygen atom

C → Carbon atom

14. Molecule

 **Definition**

A molecule is a group of two or more atoms held together by chemical bonds and capable of independent existence in the examples studied.

Examples:

Molecules of elements

H_2

O_2

Cl_2

Molecules of compounds

H_2O

HCl

CO_2

The chapter distinguishes atoms, molecules and chemical bonds before introducing covalent and ionic compounds.

15. Chemical Bond

A **chemical bond** is an attractive interaction that holds atoms or ions together.

Atoms can combine by:



Sharing electrons

or



Transferring electrons

This leads to different types of bonding.



16. Valency

Definition

Valency is the combining capacity of an atom.

It is related to the number of electrons an atom gains, loses or shares to achieve a stable electronic configuration.

Examples:

Hydrogen → Valency 1

Oxygen → Valency 2

Nitrogen → Valency 3

Carbon → Valency 4



17. Covalent Bond

A covalent bond is formed when atoms:

Share electron pairs

This type of bonding is common between non-metal atoms.



18. Formation of H₂

Each hydrogen atom has one electron.

Each needs one more electron to complete its first shell.

Two hydrogen atoms share their electrons.

Therefore:

H—H

This is a:

Single covalent bond

One shared pair of electrons = one single bond.

19. Formation of Cl₂

Each chlorine atom has:

7 valence electrons

Each needs one more electron to complete its outer shell.

Two chlorine atoms share one pair of electrons.

Therefore:



This is a:

Single covalent bond

20. Formation of O₂

Each oxygen atom has:

6 valence electrons

Each needs two more electrons.

Therefore, two oxygen atoms share:

Two pairs of electrons

Structure:



This is a:

Double covalent bond

The chapter uses chlorine and oxygen to demonstrate single and double covalent bonds.

21. Formation of HCl

Hydrogen needs:

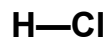
1 electron

Chlorine needs:

1 electron

They share one pair.

Therefore:



A single covalent bond is formed.

22. Formation of H₂O

Oxygen needs two more electrons.

It shares one electron with each of two hydrogen atoms.

Therefore:



Oxygen forms two covalent bonds.



23. Covalent Bond Summary

Molecule	Bond
H ₂	Single
Cl ₂	Single
HCl	Single
O ₂	Double
H ₂ O	Two O–H single bonds

★ Remember:

1 shared pair → Single bond

2 shared pairs → Double bond



24. Covalent Compounds

Compounds formed mainly through covalent bonding are called:

Covalent compounds

Examples:



25. Naming Covalent Compounds

Numerical prefixes are used to indicate the number of atoms.

Prefixes:

Number	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-

The chapter introduces these prefixes for naming covalent compounds.



26. Examples of Covalent Names



Carbon monoxide



Carbon dioxide



Carbon disulfide



Phosphorus trichloride



Sulfur hexafluoride



Dinitrogen tetroxide



Dinitrogen pentoxide



27. Reading a Chemical Formula

A chemical formula tells us:

- 1 Which elements are present
- 2 How many atoms of each element are present

Example:



H → Hydrogen

O → Oxygen

₂ → Two hydrogen atoms

Therefore:

H₂O contains 2 hydrogen atoms and 1 oxygen atom.



28. Ionic Bond

An ionic bond is formed through:

Transfer of electrons

One atom loses electrons.

Another atom gains electrons.

This produces:

+ Positive ion

and

- Negative ion

The opposite charges attract each other.



29. Formation of Sodium Chloride

Sodium electronic configuration:

2, 8, 1

Sodium has one electron in its outer shell.

It loses one electron:



Chlorine:

2, 8, 7

Chlorine gains one electron:



Now:

Na^+ attracts Cl^- .

Therefore:

NaCl

is formed.

The chapter explains NaCl formation through electron transfer and electrostatic attraction between oppositely charged ions.

30. Cation

A positively charged ion is called a:

Cation

It is generally formed by:

Loss of electrons

Examples:



31. Anion

A negatively charged ion is called an:

Anion

It is generally formed by:

Gain of electrons

Examples:



F^-

O^{2-}

S^{2-}



32. Cation vs Anion

Cation $+$

Anion $-$

Positive ion

Negative ion

Usually formed by loss of electrons

Usually formed by gain of electrons

Na^+

Cl^-

Ca^{2+}

O^{2-}

Al^{3+}

S^{2-}



Memory Trick

Cation $\rightarrow$ "+"

Anion $\rightarrow$ "-"



33. Ionic Compounds and Crystal Lattices

Ionic compounds generally do not exist as separate molecules.

Instead, their ions form:

Three-dimensional crystal lattices

For example, in sodium chloride:

Na^+ and Cl^- ions are arranged in a regular repeating structure.

The chapter notes that in the NaCl structure described, each Na^+ is surrounded by six Cl^- ions and vice versa.



34. Common Ions

+ Common Cations

+1



+2



+3



— 35. Common Anions

-1



-2



36. Important Polyatomic Ions

A polyatomic ion contains:

Two or more atoms carrying an overall charge.

Examples:

Hydroxide



Nitrate



Hydrogencarbonate



Carbonate



Sulfate



Ammonium



37. Writing Chemical Formulae

Chemical formulae show the simplest whole-number ratio of atoms or ions.

For ionic compounds:

Step 1

Write the positive ion first.

Step 2

Write the negative ion.

Step 3

Write their charges.

Step 4

Balance the charges.

Step 5

Use the smallest whole-number ratio.

Step 6

Use brackets when a polyatomic ion occurs more than once.



38. Formula of Calcium Chloride

Calcium:



Chloride:

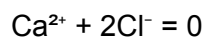


One Ca^{2+} needs two Cl^{-} ions.

Therefore:



Charge:



39. Formula of Aluminium Oxide

Aluminium:

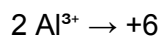


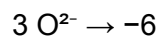
Oxide:



The smallest common charge is 6.

Therefore:





Formula:



40. Formula of Magnesium Oxide

Magnesium:



Oxide:



Charges already balance in a 1:1 ratio.

Therefore:



Not Mg_2O_2

The formula must be reduced to the simplest whole-number ratio.



41. Formula of Magnesium Hydroxide

Magnesium:



Hydroxide:



Two hydroxide ions are required.

Therefore:



⚠ **Why brackets?**

Because the entire polyatomic ion OH⁻ occurs twice.



42. Formula of Calcium Carbonate

Calcium:



Carbonate:



Charges balance in a 1:1 ratio.

Therefore:



43. Formula of Aluminium Sulfate

Aluminium:

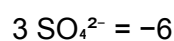
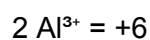


Sulfate:



LCM of 3 and 2 = 6.

Therefore:



Formula:



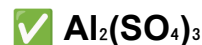
44. Common Formula-Writing Mistakes



Because the ratio 2:2 must be simplified.



Because OH is a polyatomic ion.



Brackets are needed when more than one sulfate ion is present.

45. Ionic vs Covalent Compounds

Ionic Compounds

Electron transfer

Usually metal + non-metal

Form ions

Covalent Compounds

Electron sharing

Usually non-metal + non-metal

Form molecules

Crystal lattice

Molecular structures in many cases

Generally high melting points

Generally lower melting points

Conduct when molten/aqueous

Generally poor conductors



46. Properties of Ionic Compounds

Ionic compounds generally:



Have high melting points

Strong electrostatic attractions hold the ions together.



Have high boiling points

More energy is needed to separate the ions.



Are often soluble in water

Many ionic compounds dissolve in water.



Conduct electricity when molten or dissolved

Their ions are free to move.



Do not generally conduct electricity as solids

The ions are fixed in their crystal lattice.

The chapter specifically explains that solid ionic compounds have fixed ions, while molten or aqueous ionic compounds can conduct because ions become mobile.



47. Why Does Solid NaCl Not Conduct Electricity?

In solid NaCl:

Na^+ and Cl^- ions are fixed in the crystal lattice.

They cannot move freely.

Therefore:

Solid NaCl does not conduct electricity.

48. Why Does Molten NaCl Conduct Electricity?

When NaCl is melted:

 Crystal structure breaks down.

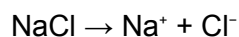
The ions become mobile.

Therefore:

Molten NaCl conducts electricity.

49. Why Does Aqueous NaCl Conduct Electricity?

When NaCl dissolves in water:



The ions can move through the solution.

Therefore:

Aqueous NaCl conducts electricity.

50. Why Does Sugar Solution Generally Not Conduct?

Sugar dissolves in water but does not normally produce ions.

Therefore:

Sugar solution generally does not conduct electricity.

This is an important comparison in the chapter.

51. Properties of Covalent Compounds

Many covalent compounds:

 **Have relatively low melting points**

 **Have relatively low boiling points**

 **Do not conduct electricity**

because they generally do not contain freely moving ions.

 **Solubility**

Many are poorly soluble in water, though there are exceptions.

Many dissolve better in organic solvents.

52. Ionic vs Covalent — Quick Comparison

Property	Ionic	Covalent
----------	-------	----------

Bond formation	Electron transfer	Electron sharing
Particles	Ions	Molecules
Structure	Crystal lattice	Molecular structures in many cases
Melting point	Generally high	Generally low
Boiling point	Generally high	Generally low
Electrical conduction	Molten/aqueous	Generally poor
Examples	NaCl, MgO	H ₂ O, CO ₂



53. Molecular Mass

Molecular mass is the:

Sum of the atomic masses of all atoms present in one molecule.

Formula:

Molecular Mass = $\Sigma(\text{Atomic mass} \times \text{Number of atoms})$

The unit commonly used is:

u



54. Molecular Mass of H₂O

Atomic masses:

H = 1 u

O = 16 u

Formula:

H₂O

Therefore:

$$\text{Molecular mass} = (2 \times 1) + 16$$

$$= 2 + 16$$

$$= 18 \text{ u}$$



55. Molecular Mass of CO_2

$$\text{C} = 12 \text{ u}$$

$$\text{O} = 16 \text{ u}$$

Formula:



$$\text{Molecular mass} = 12 + (2 \times 16)$$

$$= 12 + 32$$

$$= 44 \text{ u}$$



56. Molecular Mass of HNO_3

$$\text{H} = 1 \text{ u}$$

$$\text{N} = 14 \text{ u}$$

$$\text{O} = 16 \text{ u}$$

Formula:



$$\text{Molecular mass} = 1 + 14 + (3 \times 16)$$

$$= 1 + 14 + 48$$

$$= 63 \text{ u}$$



57. Molecular Mass of CH_4

$$\text{C} = 12 \text{ u}$$

$$\text{H} = 1 \text{ u}$$

Formula:



$$\text{Molecular mass} = 12 + (4 \times 1)$$

$$= 16 \text{ u}$$



58. Formula Unit Mass

Ionic compounds form crystal lattices rather than separate molecules.

Therefore, we use:

Formula unit mass



Definition

Formula unit mass is the sum of the atomic masses represented in the simplest ionic formula unit.



59. Formula Unit Mass of Na_2O

$$\text{Na} = 23 \text{ u}$$

$$\text{O} = 16 \text{ u}$$

Formula:



Formula unit mass = $(2 \times 23) + 16$

= $46 + 16$

= **62 u**



60. Formula Unit Mass of KCl

K = 39 u

Cl = 35.5 u

Formula unit mass = $39 + 35.5$

= **74.5 u**



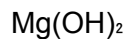
61. Formula Unit Mass of $\text{Mg}(\text{OH})_2$

Mg = 24 u

O = 16 u

H = 1 u

Formula:



Formula unit mass = $24 + 2(16 + 1)$

= $24 + 34$

= **58 u**



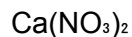
62. Formula Unit Mass of $\text{Ca}(\text{NO}_3)_2$

Ca = 40 u

N = 14 u

O = 16 u

Formula:



First calculate:

$$\text{NO}_3 = 14 + (3 \times 16)$$

$$= 14 + 48$$

$$= 62 \text{ u}$$

There are two nitrate ions:

$$2 \times 62 = 124 \text{ u}$$

Therefore:

$$\text{Formula unit mass} = 40 + 124$$

$$= 164 \text{ u}$$



63. Molecular Mass vs Formula Unit Mass

Molecular Mass	Formula Unit Mass
Used for molecules	Used for ionic compounds
Sum of atomic masses in one molecule	Sum of atomic masses in one formula unit
Example: H ₂ O	Example: NaCl
Unit: u	Unit: u



64. Important Applications

Atomic bonding and chemical compounds have applications in:



Medicine



Energy



Chemical industries



Scientific research



Agriculture

The chapter also connects atomic science with peaceful applications and highlights Indian contributions to peaceful atomic energy.



65. Scientific Thinking in This Chapter

Science is not only about memorising formulas.

The chapter encourages students to use:



Observations



Measurements



Hypotheses



Experiments



Evidence



Controlled variables

to reach scientific conclusions.



66. Important Definitions

Atom

Smallest unit of an element that retains its chemical identity.

Molecule

A group of two or more atoms held together by chemical bonds and capable of independent existence in the examples studied.

Chemical bond

Attractive interaction holding atoms or ions together.

Valency

Combining capacity of an atom.

Cation

Positively charged ion.

Anion

Negatively charged ion.

Ionic bond

Bond formed through electron transfer and electrostatic attraction between oppositely charged ions.

Covalent bond

Bond formed by sharing electron pairs.

Molecular mass

Sum of atomic masses of atoms in one molecule.

Formula unit mass

Sum of atomic masses represented in the simplest ionic formula unit.



67. Important Exam Questions

Q1. State the Law of Conservation of Mass.

Mass can neither be created nor destroyed in a chemical reaction.

Total mass of reactants = Total mass of products.

Q2. State the Law of Constant Proportions.

A pure compound always contains the same elements combined in the same fixed ratio by mass.

Q3. State the main ideas of Dalton's atomic theory.

Matter consists of atoms; atoms combine in simple whole-number ratios; and chemical reactions involve rearrangement of atoms.

Q4. What is a molecule?

A molecule is a group of two or more atoms held together by chemical bonds and capable of independent existence in the examples studied.

Q5. What is a covalent bond?

A covalent bond is formed by sharing electron pairs between atoms.

Q6. What is an ionic bond?

An ionic bond results from electron transfer followed by electrostatic attraction between oppositely charged ions.

Q7. What is a cation?

A positively charged ion.

Q8. What is an anion?

A negatively charged ion.

Q9. Why do ionic compounds conduct electricity when molten?

Because their ions become free to move.

Q10. Why does solid NaCl not conduct electricity?

Because its ions are fixed in the crystal lattice.

Q11. Why does aqueous NaCl conduct electricity?

Because NaCl produces mobile Na^+ and Cl^- ions in water.

Q12. Why does sugar solution generally not conduct electricity?

Because sugar dissolves without producing ions.

Q13. What is molecular mass?

The sum of the atomic masses of all atoms present in one molecule.

Q14. What is formula unit mass?

The sum of the atomic masses represented in the simplest ionic formula unit.

68. Competency-Based Question

A student mixes two substances and finds that their ratio changes depending on the amounts used.

Is the resulting material necessarily a pure compound?

 **No.**

A pure compound has a fixed ratio of its elements.

A mixture can have variable proportions.

69. Competency-Based Question

A chemical reaction takes place in an open flask and produces a gas.

The mass measured after the reaction is less than before.

Has the law of conservation of mass been violated?

 **No.**

The gas has escaped from the open system.

If the entire system, including the escaped gas, were accounted for:

Total mass would be conserved.

70. Competency-Based Question

An element has electronic configuration:

2, 8, 1

What is likely to happen to its outer electron?

It can lose one electron to achieve a more stable configuration.

It forms:

A +1 cation

Example:



71. Competency-Based Question

An element has configuration:

2, 8, 7

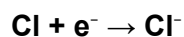
What is it likely to do?

It can gain one electron to complete its outer shell.

It forms:

A -1 anion

Example:



72. Competency-Based Question

Why is the formula of aluminium oxide Al_2O_3 ?

Aluminium:



Oxygen:



Two Al^{3+} ions give:

+6

Three O^{2-} ions give:

-6

Therefore, total charge becomes zero.

Formula = Al_2O_3

73. Competency-Based Question

Why is magnesium hydroxide written as $\text{Mg}(\text{OH})_2$ and not MgOH_2 ?

Because hydroxide:



is a polyatomic ion.

Two hydroxide ions are needed.

Therefore, brackets are required:



74. Common Mistakes to Avoid

Mistake 1

Saying mass is destroyed when gas escapes.

Correct

Matter has escaped from the open system.

Mistake 2

Calling a compound a mixture.

Correct

A compound has a fixed composition.

A mixture can have variable composition.

Mistake 3

Saying ionic compounds contain separate molecules.

Correct

Ionic compounds generally form crystal lattices.

Mistake 4

Confusing ionic and covalent bonding.

Remember:

Ionic → **Transfer**

Covalent → **Sharing**

Mistake 5

Writing Mg_2O_2 instead of MgO .

Correct:

Always reduce the ionic formula to the simplest whole-number ratio.

✗ Mistake 6

Writing AlOH_3 .

✓ **Correct:**



Use brackets around a repeated polyatomic ion.

✗ Mistake 7

Putting charges into the final neutral formula.

✓ **Correct:**

Charges are used to determine the ratio but are not written in the final neutral formula.

75. QUICK REVISION

- ★ Law of Conservation of Mass → Total mass remains constant.
- ★ Law of Constant Proportions → Pure compounds have fixed mass ratios.
- ★ Dalton → Matter is made of atoms.
- ★ Chemical reactions → Rearrangement of atoms.
- ★ Atom → Basic unit of an element.
- ★ Molecule → Group of bonded atoms.
- ★ Valency → Combining capacity.
- ★ Covalent bond → Sharing of electrons.
- ★ Ionic bond → Transfer of electrons + attraction.
- ★ Cation → Positive ion.
- ★ Anion → Negative ion.

- ★ Ionic compounds → Crystal lattices.
 - ★ Solid ionic compounds → Generally do not conduct.
 - ★ Molten/aqueous ionic compounds → Conduct because ions can move.
 - ★ Covalent compounds → Generally poor electrical conductors.
 - ★ Molecular mass → Sum of atomic masses in a molecule.
 - ★ Formula unit mass → Sum of atomic masses in an ionic formula unit.
-

76. MEMORY TRICKS

CONSERVATION

Reactants Mass = Products Mass

CONSTANT PROPORTIONS

Compound → Fixed Ratio

DALTON

Atoms → Combine → Rearrange

COVALENT

Covalent = Co-share

 **Sharing electrons**

IONIC

Ionic = Ions

👉 Transfer electrons

+ CATION

Cation = Positive

— ANION

Anion = Negative

📊 MASS

Molecular Mass → Molecule

Formula Unit Mass → Ionic Compound



77. COMPLETE FORMULA SHEET

Law of Conservation of Mass

Mass of reactants = Mass of products

Mass of product

Mass of product = Total mass of reactants – Mass of other products

Constant Proportion

Mass ratio of elements in a pure compound = Constant

Number of atoms

The subscript in a chemical formula tells the number of atoms of that element.

Molecular Mass

$\text{Molecular Mass} = \Sigma(\text{Atomic mass} \times \text{Number of atoms})$

Formula Unit Mass

$\text{Formula Unit Mass} = \Sigma(\text{Atomic mass} \times \text{Number of atoms in formula unit})$

Charge Balance

$\text{Total positive charge} = \text{Total negative charge}$

for a neutral ionic compound.

78. CHAPTER 9 — FINAL TAKEAWAY

This chapter explains **why elements combine in definite ways**. 

The journey is:

 **Laws of Chemical Combination**



 **Atoms**



 **Molecules**



 **Chemical Bonds**



+ — Ions



 **Chemical Formulae**



Molecular Mass / Formula Unit Mass

The three biggest ideas to remember are:

★ **Conservation of Mass**

Total mass of reactants = Total mass of products

★ **Constant Proportions**

A pure compound always has a fixed ratio of elements by mass

★ **Chemical Bonding**

Covalent = Sharing | Ionic = Transfer

And for numerical questions, remember:

Molecular Mass = Add atomic masses in a molecule

Formula Unit Mass = Add atomic masses in the simplest ionic formula unit

The supplied chapter notes summarize the complete conceptual chain as **law** → **composition** → **particles** → **bonding** → **formula** → **mass**.