



EXPANDED STUDY MATERIAL

CHAPTER 9

Atomic Foundations of Matter

Chemical laws, atoms, molecules, bonding and formulae

INSIDE THIS BOOKLET

Detailed theory notes • Key definitions • Formula and concept boxes • Worked examples • One full page of practice questions • Answer key at the end

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Use these notes alongside your prescribed textbook and classroom teaching. Scientific terms and examples are explained in student-friendly language for revision and practice.

THEORY NOTES

1. From atoms to compounds

An element is a pure substance made of one kind of atom. A compound is a pure substance formed when two or more elements combine chemically in a fixed proportion. The properties of a compound are often very different from the properties of its constituent elements. For example, hydrogen burns and oxygen supports combustion, while water formed from them can extinguish many fires.

A mixture differs from a compound. In a mixture, substances are physically combined in variable proportions and can often be separated by physical methods. In a compound, elements are chemically combined in a definite ratio and separation generally requires chemical change.

Atoms are the basic units used to describe chemical elements. Some elements exist as single atoms under ordinary conditions, while others form molecules such as O_2 , N_2 or Cl_2 . A molecule is a group of atoms bonded together and acting as a unit. Compounds may contain molecules (such as H_2O) or ions arranged in a crystal lattice (such as sodium chloride).

2. Laws of chemical combination

The **law of conservation of mass** states that mass is neither created nor destroyed in a chemical reaction. In a closed system, the total mass of reactants equals the total mass of products. If a reaction appears to lose mass in an open container, a gas may have escaped; if it appears to gain mass, a gas from the air may have entered.

The **law of definite proportions** (constant composition) states that a pure chemical compound always contains the same elements combined in the same fixed proportion by mass, regardless of its source or method of preparation. Pure water, for example, always has hydrogen and oxygen in the same mass ratio.

The **law of multiple proportions** applies when two elements form more than one compound. For a fixed mass of one element, the masses of the other element that combine with it are in simple whole-number ratios. This pattern supports the idea that matter combines in discrete units.

These laws describe regularities observed in chemical reactions. They are not just memorised statements: they can be checked by measuring reactant and product masses carefully, ensuring gases are included and experimental losses are controlled.

Chemistry reminder

A balanced chemical formula is electrically neutral overall; use the smallest whole-number ratio of ions.

3. Dalton's atomic theory and its development

Dalton's atomic theory proposed that matter is made of atoms; atoms of an element were treated as identical in mass and properties; atoms of different elements were considered different; compounds form when atoms combine in simple whole-number ratios; and chemical reactions rearrange atoms rather than create or destroy them.

Modern discoveries refined parts of Dalton's theory. Atoms contain electrons, protons and neutrons, so they are divisible. Atoms of one element can have different masses because of isotopes. However, the central idea that chemical reactions rearrange atoms and that compounds have definite composition remains fundamental.

Relative atomic mass compares the average mass of atoms of an element with a standard reference. In school calculations, atomic masses from the periodic table are used to calculate formula or molecular masses. The unit u is used for atomic-scale masses.

4. Valency and chemical bonding

Atoms form chemical bonds because the bonded arrangement can be more stable than separate atoms. For many main-group elements, stability is associated with a filled outer electron shell. Atoms may lose, gain or share valence electrons. The outermost-shell electrons involved in bonding are called valence electrons.

Ionic bonding commonly involves electron transfer. A metal atom may lose electrons to form a positive ion (cation), while a non-metal gains electrons to form a negative ion (anion). Oppositely charged ions attract electrostatically. Sodium chloride contains Na^+ and Cl^- ions in a repeating lattice; it is not made of separate $NaCl$ molecules in the solid crystal.

Covalent bonding involves sharing pairs of electrons between atoms. Water has covalent bonds between oxygen and hydrogen; oxygen molecules contain a double bond in the usual school-level model. Covalent substances can exist as molecules or giant networks.

Valency is a simple measure of combining capacity. It is often found by considering how many electrons an atom loses, gains or shares to reach a stable outer shell. Common valencies include $H = 1$, $O = 2$, $N = 3$, $C = 4$, $Na = 1$, $Mg = 2$, $Al = 3$ and $Cl = 1$. Some elements show more than one valency in different compounds.

5. Writing chemical formulae

A chemical formula uses element symbols and subscripts to show the types and relative numbers of atoms or ions in a substance. H_2O means two hydrogen atoms for each oxygen atom in a water molecule; CO_2 means one carbon atom and two oxygen atoms.

To write a formula from valencies, write the positive part first and the negative part second, then balance the combining capacities using the smallest whole-number ratio. For magnesium chloride, Mg has valency 2 and Cl has valency 1, so the formula is MgCl_2 . For aluminium oxide, Al has valency 3 and O has valency 2, giving Al_2O_3 .

Polyatomic ions such as sulfate (SO_4^{2-}), nitrate (NO_3^-), hydroxide (OH^-) and ammonium (NH_4^+) act as units. If more than one polyatomic ion is required, use brackets: $\text{Ca}(\text{OH})_2$. Do not change the internal subscripts of a polyatomic ion when balancing a formula.

A formula represents the simplest ratio for ionic compounds and the actual atom count for a molecule. Subscripts are not written as 1, and the charges on ions are not normally written in the final neutral compound formula.

6. Molecular and formula mass

Molecular mass is the sum of the relative atomic masses of all atoms in a molecule. For H_2O , using $\text{H} = 1$ and $\text{O} = 16$, molecular mass = $2(1) + 16 = 18 \text{ u}$. For CO_2 , it is $12 + 2(16) = 44 \text{ u}$.

For ionic compounds, the more precise term is formula mass, because the solid is an extended arrangement of ions rather than separate molecules. The calculation is still the sum of atomic masses in the formula unit. For NaCl , formula mass $\approx 23 + 35.5 = 58.5 \text{ u}$.

Brackets and subscripts must be handled carefully. $\text{Ca}(\text{OH})_2$ contains one Ca, two O and two H atoms, so its formula mass using $\text{Ca} = 40$, $\text{O} = 16$ and $\text{H} = 1$ is $40 + 2(16 + 1) = 74 \text{ u}$.

7. Solved examples

Example 1 — Conservation of mass: If 12 g of carbon reacts completely with 32 g of oxygen, the mass of carbon dioxide formed is $12 + 32 = 44 \text{ g}$, assuming no material escapes.

Example 2 — Formula: Aluminium has valency 3 and oxygen valency 2. Balance the valencies to obtain Al_2O_3 .

Example 3 — Molecular mass: $\text{H}_2\text{SO}_4 = 2(1) + 32 + 4(16) = 98 \text{ u}$.

Example 4 — Formula mass: $\text{Ca}(\text{OH})_2 = 40 + 2(16 + 1) = 74 \text{ u}$.

QUICK REVISION SUMMARY

Key ideas to retain

Chemical reactions rearrange atoms. Mass is conserved in a closed system. Compounds have fixed composition. Valency helps write formulae. Molecular/formula mass is the sum of the atomic masses represented by the formula.

PRACTICE QUESTIONS

Attempt all questions before checking the answer key. Show working for numerical questions and write full sentences for explanation questions.

1. Distinguish between an element, a compound and a mixture.
2. State the law of conservation of mass.
3. Explain the law of definite proportions.
4. State one important limitation of Dalton's original atomic theory.
5. What is the difference between ionic and covalent bonding?
6. Write the formula for magnesium oxide using Mg valency 2 and O valency 2.
7. Write the formula for aluminium chloride (Al valency 3, Cl valency 1).
8. Write the formula for calcium hydroxide using Ca^{2+} and OH^- .
9. Calculate the molecular mass of NH_3 (N = 14, H = 1).
10. Calculate the molecular mass of CO_2 (C = 12, O = 16).
11. Calculate the formula mass of Na_2CO_3 (Na = 23, C = 12, O = 16).
12. A reaction starts with 18 g of reactants and produces 11 g of one product. What mass of the other product is formed if the system is closed?
13. Explain why a solid ionic compound is better described using formula units than separate molecules.
14. What do the subscripts in H_2SO_4 tell you?
15. Two elements form compounds AB and A_2B_3 . What does the existence of multiple compounds illustrate?

Before you submit

Check calculations, include units where appropriate, define important terms clearly and support explanations with examples.

ANSWER KEY

Use the answers to check your understanding. For explanation questions, equivalent scientifically correct wording is acceptable.

1. Element: one kind of atom; compound: elements chemically combined in a fixed ratio; mixture: substances physically combined in variable proportions.
2. Total mass of reactants equals total mass of products in a closed system.
3. A pure compound always contains the same elements in the same fixed ratio by mass.
4. Atoms are divisible into subatomic particles, and isotopes of one element can have different masses.
5. Ionic bonding involves attraction between oppositely charged ions after electron transfer; covalent bonding involves shared electron pairs.
6. MgO.
7. AlCl_3 .
8. Ca(OH)_2 .
9. $14 + 3(1) = 17 \text{ u}$.
10. $12 + 2(16) = 44 \text{ u}$.
11. $2(23) + 12 + 3(16) = 106 \text{ u}$.
12. Total products must be 18 g; other product = $18 - 11 = 7 \text{ g}$.
13. Ions form a repeating crystal lattice rather than discrete molecules; the formula gives the simplest ion ratio.
14. Each formula unit has 2 H atoms, 1 S atom and 4 O atoms.
15. It illustrates that the same elements can form more than one compound in different whole-number ratios.

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Revise the theory, practise without looking at the answers, and revisit any question you could not explain in your own words.

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