



Counting the Uncountable

Atoms, Molecules and the magical number called the Mole

INSTITUTE Prodigy Educlasses	CLASS IX (Class 9)	BOARD CBSE / ICSE / WB Board
TOPIC Atoms and Molecules	CHAPTER / UNIT Atoms and Molecules (Mole Concept)	FACULTY Surajit Sir (Chemistry)

Let me begin with a number that will blow your mind.

A single drop of water holds about 5×10^{21} atoms. That is 5 followed by 21 zeros. If you counted one atom every second, without ever sleeping, it would take you roughly **160 trillion years**. The universe itself is only about 13.8 billion years old, so you would need more than 11,000 universes' worth of time. And yet chemists count atoms every day. By the end of this note, you will know their secret.



The Mystery Box: 4 questions to keep in your head

- A candle burns and seems to 'vanish'. Where does its mass go?
- Water from a river and water made in a lab always have hydrogen and oxygen in the same ratio. Why?
- Nobody can weigh one atom on a kitchen scale. So how do we count atoms at all?
- What is a 'mole', and why does it have nothing to do with the animal?

1 Two laws that chemistry stands on

In the 1700s, chemists were busy weighing everything before and after reactions. Two patterns kept appearing so reliably that they became laws.

Law 1: Conservation of mass (Lavoisier). In a chemical reaction, mass is neither created nor destroyed. Total mass of reactants = total mass of products.

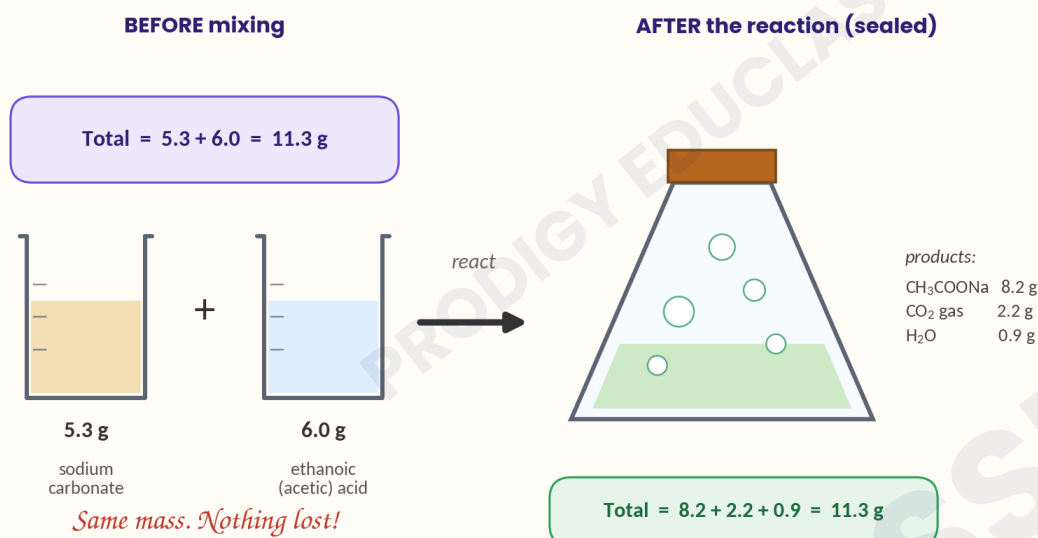


Fig 1. Sodium carbonate + ethanoic acid in a sealed flask. The total mass stays exactly the same.

⚠ Careful! Common Mistake

'But when I burn paper, the ash weighs less!' That is because the gases (carbon dioxide and water vapour) escape into the air. Burn the paper in a sealed container and weigh everything, and the mass matches perfectly. The mass never vanished. It just left the scale.

Law 2: Constant proportions (Proust). A pure chemical compound always contains the same elements combined in the same fixed ratio by mass, no matter where it comes from or how it is made.

Compound	Elements	Ratio by mass	What it means
Water, H ₂ O	Hydrogen : Oxygen	1 : 8	9 g of water always holds 1 g H and 8 g O
Carbon dioxide, CO ₂	Carbon : Oxygen	3 : 8	44 g always holds 12 g C and 32 g O
Ammonia, NH ₃	Nitrogen : Hydrogen	14 : 3	17 g holds 14 g N and 3 g H
Methane, CH ₄	Carbon : Hydrogen	3 : 1	16 g holds 12 g C and 4 g H

🔑 Teacher's Trick

Quick check. 3 g of hydrogen burns completely with oxygen. How much oxygen is needed, and how much water forms? Ratio is 1 : 8, so oxygen = $3 \times 8 = 24$ g, and water = $3 + 24 = 27$ g. Both laws in one problem!

2 Dalton's big idea: everything is made of atoms

How can we explain the two laws? In 1808, the English teacher and chemist John Dalton had a beautiful answer. Matter is made of tiny particles called atoms, and chemical reactions just rearrange them. Nothing is created or destroyed, so mass is conserved. And if a compound always has the same atoms in the same numbers, the ratio of masses must stay fixed too.

Dalton said (1808)	Today we know
All matter is made of tiny indivisible particles called atoms	Atoms can be split into protons, neutrons and electrons (you will meet them next chapter)
Atoms of the same element are identical in mass and properties	Mostly true, but isotopes of an element have slightly different masses
Atoms of different elements differ in mass and properties	Correct
Atoms combine in small whole-number ratios to form compounds	Correct. This is why H_2O is always H_2O
Atoms cannot be created or destroyed in a reaction	Correct for chemical reactions (nuclear reactions are a different story)

Did you know?

The idea is older than Dalton. Around 2,500 years ago the Indian sage Kanada wrote about *parmanu*, tiny indivisible particles of matter. In Greece, Democritus called them *atomos*, meaning 'uncuttable'. What Dalton added was the measuring: he connected atoms to real weights and real experiments. And notice that he was not 100% right. Science is not about being right the first time. It is about getting closer.

3 Atoms, molecules and ions: meet the particles

- **Atom:** the smallest particle of an element that can take part in a chemical reaction. Example: one atom of helium, He.
- **Molecule:** two or more atoms held together. It can be an element (O_2 , N_2 , P_4 , S_8) or a compound (H_2O , CO_2).
- **Ion:** an atom or group of atoms with an electric charge. Positive = cation (Na^+), negative = anion (Cl^-). Ionic compounds like NaCl form endless crystal lattices, so we write the simplest ratio instead of a molecule.
- **Atomicity:** the number of atoms in a molecule. He is monatomic, O_2 diatomic, O_3 and H_2O triatomic, P_4 tetratomic.

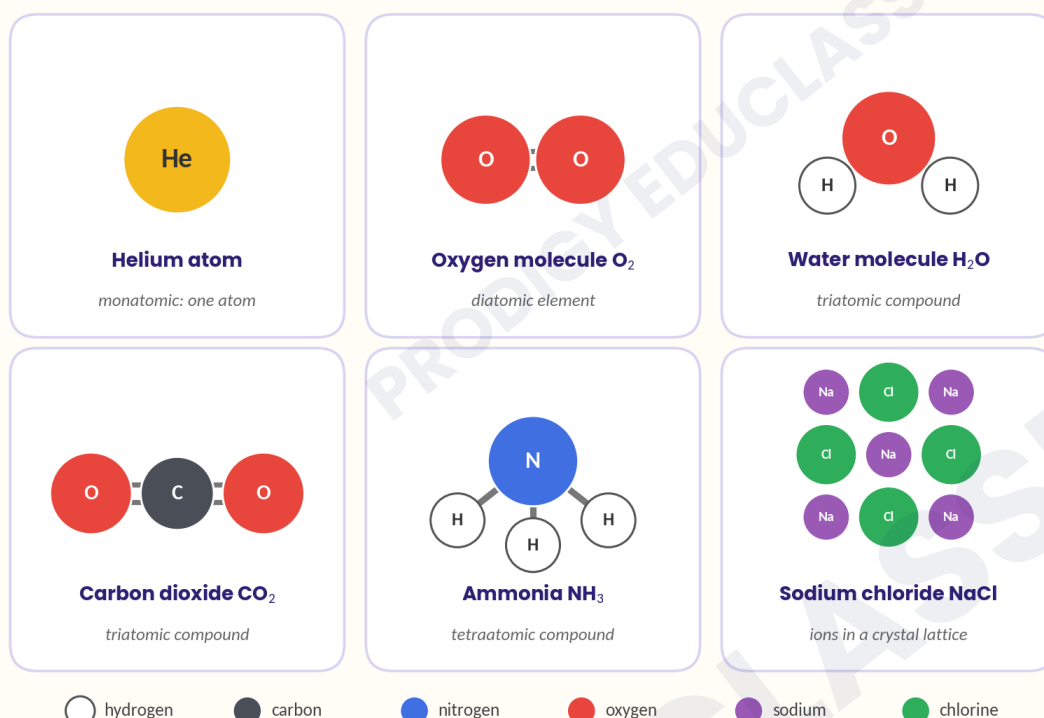


Fig 2. A gallery of particles. Notice that $NaCl$ has no separate 'molecules', only a repeating pattern of ions.

How small is an atom, really? Figure 3 places it on a ruler where every tick is ten times bigger than the last one.

Each step to the right is 10 times bigger

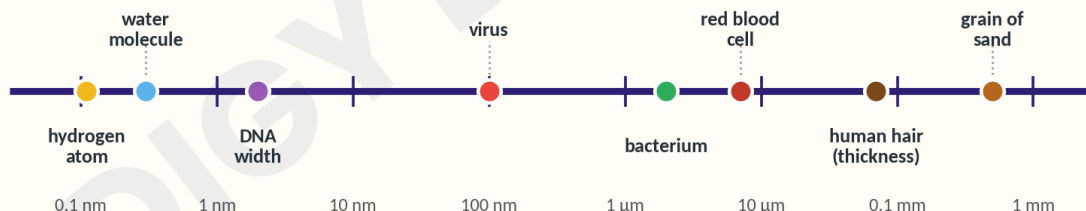


Fig 3. From a hydrogen atom to a grain of sand: a 5 million-fold journey (values approximate).

Why is sodium 'Na' and not 'So'? Some symbols come from old Latin names. Here are the ones you must know:

Symbol	Comes from (Latin)	Element	Symbol	Comes from (Latin)	Element
Na	natrium	Sodium	Au	aurum	Gold
K	kalium	Potassium	Ag	argentum	Silver
Fe	ferrum	Iron	Pb	plumbum	Lead
Cu	cuprum	Copper	Hg	hydrargyrum	Mercury

4 Weighing the invisible: atomic and molecular mass

We cannot use grams for a single atom: the numbers would be absurdly tiny (one carbon atom weighs about 2×10^{-23} g). So scientists made a new unit, the **atomic mass unit (u)**. One u is exactly **1/12 the mass of one carbon-12 atom**. On this scale: H = 1 u, C = 12 u, N = 14 u, O = 16 u, Na = 23 u, Mg = 24 u, Al = 27 u, S = 32 u, Cl = 35.5 u, K = 39 u, Ca = 40 u.

The **molecular mass** of a substance is just the sum of the atomic masses of all atoms in its formula. Do it once and you will never forget it:

Substance	Working	Mass
Water, H_2O	$2 \times 1 + 16$	18 u
Carbon dioxide, CO_2	$12 + 2 \times 16$	44 u
Ammonia, NH_3	$14 + 3 \times 1$	17 u
Sulphuric acid, H_2SO_4	$2 \times 1 + 32 + 4 \times 16$	98 u
Calcium carbonate, CaCO_3	$40 + 12 + 3 \times 16$	100 u
Glucose, $\text{C}_6\text{H}_{12}\text{O}_6$	$6 \times 12 + 12 \times 1 + 6 \times 16$	180 u
Common salt, NaCl (formula mass)	$23 + 35.5$	58.5 u

5 Writing formulae in 30 seconds: the criss-cross trick

Every ion has a combining power called **valency** (its charge, ignoring the sign). To write a formula, swap the valencies and drop them as subscripts. That is it. First learn the common ions:

Valency	Positive ions (cations)	Negative ions (anions)
1	Na^+ , K^+ , H^+ , NH_4^+ , Ag^+	Cl^- , OH^- , NO_3^- , HCO_3^-
2	Mg^{2+} , Ca^{2+} , Zn^{2+} , Cu^{2+} , Fe^{2+}	O^{2-} , S^{2-} , SO_4^{2-} , CO_3^{2-}
3	Al^{3+} , Fe^{3+}	PO_4^{3-} , N^{3-}

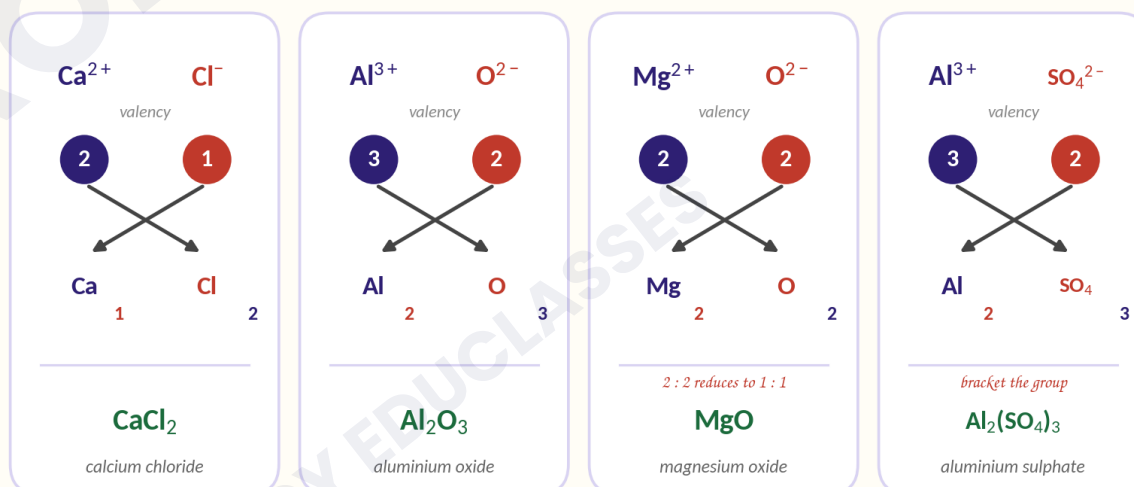


Fig 4. Swap the valencies, drop them as subscripts, then simplify. Use brackets when a group of atoms needs a number.

⚠ Careful! Common Mistake

Three slip-ups I see every year: (1) forgetting to reduce, so MgO becomes Mg_2O_2 ; (2) writing SO_43 instead of $(\text{SO}_4)_3$; (3) changing the small numbers inside a group. You can only change subscripts **outside** the bracket while writing the formula.

6 The mole: a chemist's dozen

You buy bananas by the dozen, not one by one. You buy paper by the ream (500 sheets). Chemists also need a 'counting unit'. But atoms are so tiny that even a million is a small number. So they chose an enormous group:

$$1 \text{ mole} = 6.022 \times 10^{23} \text{ particles (this is the Avogadro constant, } N_A)$$

The particles can be atoms, molecules, ions or anything else. One mole of atoms, one mole of molecules, one mole of sugar grains: always 6.022×10^{23} . Now comes the real magic, the reason chemists love the mole:

⚖ Teacher's Trick

The mass of one mole, in grams, equals the atomic or molecular mass in u. Carbon has atomic mass 12 u, so **12 g of carbon = 1 mole = 6.022×10^{23} atoms**. Water has molecular mass 18 u, so 18 g of water = 1 mole of molecules. This lets you count particles by simply weighing. That is how nobody-can-weigh-one-atom becomes possible.

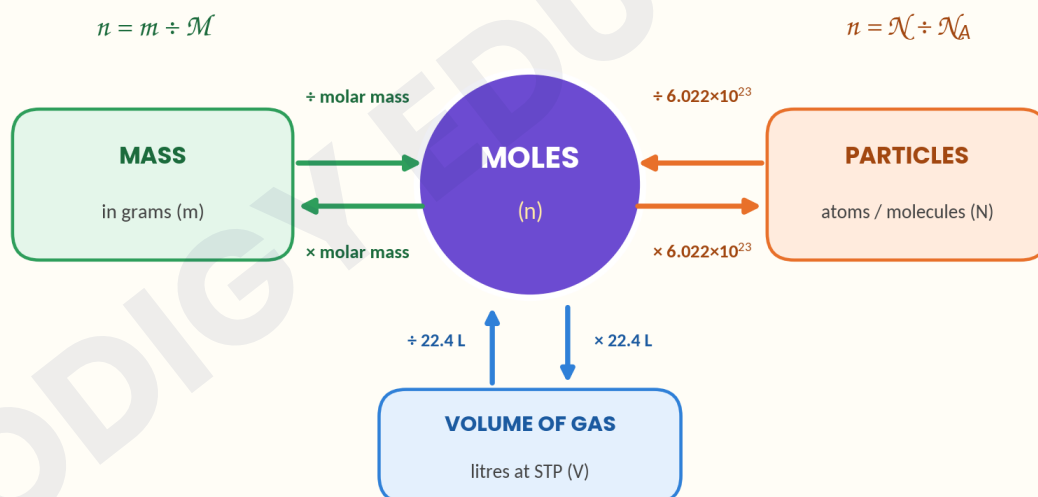


Fig 5. The mole bridge. The mole sits in the middle, and mass, particles and gas volume all connect to it.

One mole of...	Molar mass	Contains
Carbon, C	12 g	6.022×10^{23} carbon atoms
Water, H_2O	18 g (about a small sip)	6.022×10^{23} molecules
Common salt, NaCl	58.5 g	6.022×10^{23} formula units
Iron, Fe	56 g	6.022×10^{23} iron atoms
Sugar, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$	342 g	6.022×10^{23} molecules

Two formulas do almost every problem you will meet:

$$n = m \div M \text{ and } n = N \div N_A \text{ (} n = \text{moles, } m = \text{given mass, } M = \text{molar mass, } N = \text{number of particles)}$$

Worked example 1. Find the mass of 3 moles of water.

$$m = n \times M = 3 \times 18 = 54 \text{ g.}$$

Worked example 2. How many molecules are there in 88 g of carbon dioxide?

$$\text{Moles} = 88 \div 44 = 2 \text{ mol. Molecules} = 2 \times 6.022 \times 10^{23} = 1.2044 \times 10^{24}.$$

Worked example 3. How many atoms are in 9 g of water?

$$\text{Moles} = 9 \div 18 = 0.5 \text{ mol. Molecules} = 0.5 \times 6.022 \times 10^{23} = 3.011 \times 10^{23}. \text{ Each molecule has 3 atoms (2 H + 1 O), so atoms} = 3 \times 3.011 \times 10^{23} = 9.033 \times 10^{23}.$$

Worked example 4. What is the mass of a single atom of carbon?

$$12 \text{ g holds } 6.022 \times 10^{23} \text{ atoms, so one atom} = 12 \div (6.022 \times 10^{23}) \approx 1.99 \times 10^{-23} \text{ g.}$$

Exam Alert

Read the question for the word **atoms** or **molecules**. In Example 3 the answer for molecules is 3.011×10^{23} , but for atoms it is three times more. Many marks are lost right here. Also write units at every step.

Did you know?

Avogadro never knew his own number. Amedeo Avogadro (1776 to 1856), an Italian scientist, suggested in 1811 that equal volumes of gases hold equal numbers of molecules. The actual count was measured decades after his death, and the constant was named in his honour.

Did you know?

Big numbers, made friendly. A mole of seconds is about 19 quadrillion years. One glass of water (250 mL) contains roughly 8×10^{24} molecules, while all the world's oceans together hold only about 5×10^{21} glasses of water. So one glass holds more molecules than the oceans hold glasses. And your own body carries about 7×10^{27} atoms.

7 One-page cheat sheet

- ✓ **Conservation of mass:** mass of reactants = mass of products. **Constant proportions:** a compound always has fixed mass ratios (water H : O = 1 : 8).
- ✓ **Dalton:** matter is made of atoms; atoms are rearranged, not created or destroyed, in reactions.
- ✓ Atom, molecule and ion are different things. Atomicity = number of atoms in a molecule.
- ✓ $1 \text{ u} = 1/12$ of a carbon-12 atom. Molecular mass = sum of atomic masses.
- ✓ Formula writing: swap valencies, drop them, reduce, and use brackets for groups.
- ✓ **1 mole = 6.022×10^{23} particles.** Mass of 1 mole in grams = molecular mass in u.
- ✓ $n = m \div M$ and $n = N \div N_A$. At STP, 1 mole of any gas occupies 22.4 L (sneak peek for Class 11).

8 Test yourself

- Q1. 6 g of carbon burns in exactly 16 g of oxygen. What mass of carbon dioxide forms, and which law tells you?
- Q2. Write the formulae of: magnesium chloride, calcium hydroxide, sodium sulphate, aluminium nitrate.
- Q3. Calculate the molecular mass of H_2SO_4 and of CaCO_3 .
- Q4. How many moles and how many molecules are there in 36 g of water?
- Q5. What is the mass of 0.5 mol of CO_2 ?
- Q6. How many atoms are present in 0.1 mol of oxygen gas (O_2)?
- Q7. Why does burnt paper leave ash that weighs less than the paper, even though mass is conserved?
- Q8. Give two differences between an atom and a molecule, and name two diatomic elements.

Peek only after trying!

Answers. 1) 22 g of CO_2 , law of conservation of mass ($6 : 16 = 3 : 8$, the constant proportion too). 2) MgCl_2 , Ca(OH)_2 , Na_2SO_4 , $\text{Al(NO}_3)_3$. 3) 98 u and 100 u. 4) $36 \div 18 = 2$ mol; $2 \times 6.022 \times 10^{23} = 1.2044 \times 10^{24}$ molecules. 5) $0.5 \times 44 = 22$ g. 6) 0.1 mol of $\text{O}_2 = 6.022 \times 10^{22}$ molecules, and each has 2 atoms, so 1.2044×10^{23} atoms. 7) Gases (CO_2 , water vapour) escape into the air. 8) An atom is the single particle of an element that takes part in reactions, a molecule is a group of atoms bonded together. Diatomic examples: H_2 , O_2 , N_2 , Cl_2 .



A last word from your teacher

Atoms are too small to see, too many to count, and yet we can weigh them, count them and predict what they will do. That is the power of a good idea. Stay curious, and never be afraid of a big number. **Keep exploring!** - Surajit Sir