

Counting particles by mass: the mole

International Baccalaureate · IB Diploma · Chemistry · SL

Scope and prerequisites

This handout teaches Structure 1.4, Counting particles by mass: the mole, for the course first assessed in 2025. All of it is SL and HL content.

Prerequisites: atoms, isotopes and ions from Structure 1.2, chemical formulae, and standard form (6.02×10^{23}) on a calculator. Use the relative atomic masses from the data booklet.

The mole

■ How do you count grains of rice?

Nobody counts rice one grain at a time. A shop sells it by mass: if one grain has a mass of 0.02 g, then 1 kg holds about 50 000 grains. Chemists have the same problem with atoms, which are far too small and too many to count. They also count by weighing.

■ The Avogadro constant

The **mole** (symbol mol) is the unit of **amount of substance**. One mole contains exactly 6.02×10^{23} particles. This number is the **Avogadro constant**, $N_A = 6.02 \times 10^{23} \text{ mol}^{-1}$.

$$N = n \times N_A$$

N is the number of particles and n is the amount in moles.

- Always say which particles you mean: atoms, molecules, ions or **formula units**.
- One mole of H_2O molecules contains 2 mol of H atoms and 1 mol of O atoms.

■ Worked example: counting particles

How many molecules and how many atoms are in 0.50 mol of water?

- Molecules: $0.50 \times 6.02 \times 10^{23} = 3.01 \times 10^{23}$.
- Each H_2O molecule has 3 atoms, so atoms = $3 \times 3.01 \times 10^{23} = 9.03 \times 10^{23}$.

Relative masses and molar mass

■ Relative atomic mass

The **relative atomic mass** A_r is the weighted mean mass of the atoms of an element, relative to one twelfth of the mass of a carbon-12 atom. It has no units.



$$A_r = \frac{(75 \times 35) + (25 \times 37)}{100} = 35.5$$

A_r is a weighted mean of the isotope masses, so it is rarely a whole number

Natural chlorine is about 75% ^{35}Cl and 25% ^{37}Cl , so its A_r lies closer to 35 than to 37. The data booklet value is 35.45.

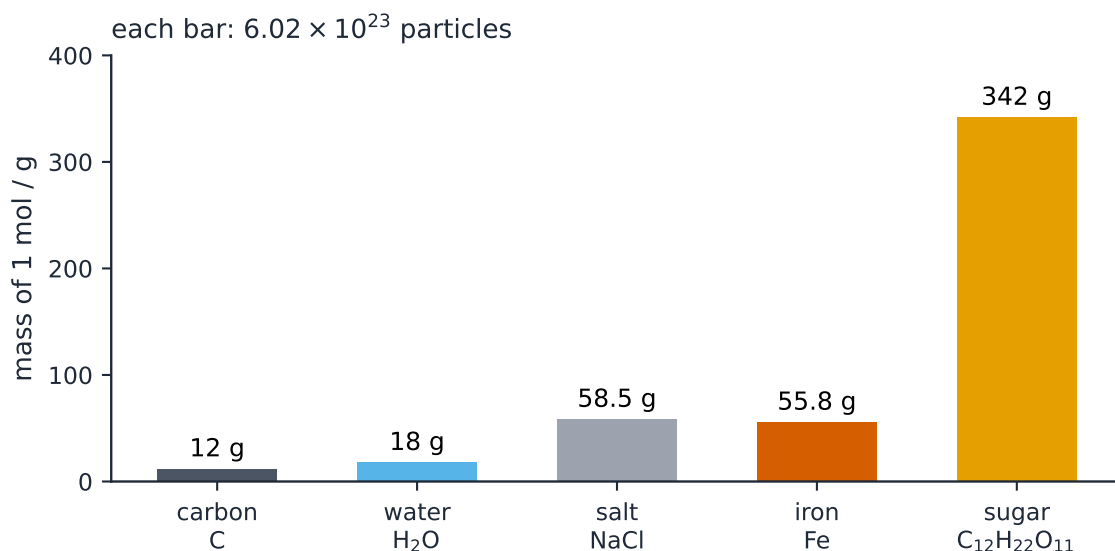
■ Relative formula mass

The **relative formula mass** M_r is the sum of the A_r values of all the atoms in the formula. "Relative molecular mass" is used for molecules.

$$M_r(\text{CaCO}_3) = 40.08 + 12.01 + 3 \times 16.00 = 100.09$$

■ Molar mass

The **molar mass** M is the mass of one mole of a substance. It has the same number as M_r , with units g mol^{-1} : $M(\text{CaCO}_3) = 100.09 \text{ g mol}^{-1}$.



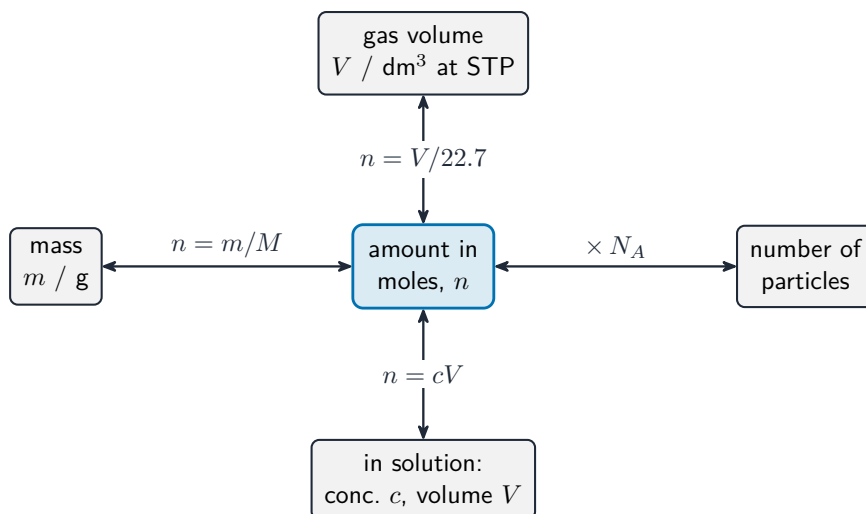
One mole of each substance has the same number of particles but a different mass

Moles, mass and the mole map

■ Mass to moles

$$n = \frac{m}{M}$$

m in g, M in g mol^{-1} , n in mol.



Every route between quantities passes through the amount in moles

■ Worked example: mass to particles

How many molecules are in 4.4 g of carbon dioxide?

- $M(\text{CO}_2) = 12.01 + 2 \times 16.00 = 44.01 \text{ g mol}^{-1}$.
- $n = 4.4/44.01 = 0.10 \text{ mol}$.
- $N = 0.10 \times 6.02 \times 10^{23} = 6.0 \times 10^{22} \text{ molecules}$.

■ Worked example: moles to mass

What is the mass of 0.250 mol of sodium chloride?

- $M(\text{NaCl}) = 22.99 + 35.45 = 58.44 \text{ g mol}^{-1}$.
- $m = n \times M = 0.250 \times 58.44 = 14.6 \text{ g}$.

■ Percentage by mass

$$\% \text{ of element} = \frac{(\text{number of atoms}) \times A_r}{M_r} \times 100$$

For ammonium nitrate, NH_4NO_3 : $M_r = 80.06$, and it has 2 N atoms, so $\% \text{ N} = 2 \times 14.01/80.06 \times 100 = 35.0\%$. Farmers compare fertilizers this way.

Empirical and molecular formulae

■ Two kinds of formula

- The **empirical formula** is the simplest whole-number ratio of the atoms of each element in a compound.
- The **molecular formula** is the actual number of atoms of each element in one molecule.

Compound	Molecular formula	Empirical formula
ethane	C_2H_6	CH_3
glucose	$\text{C}_6\text{H}_{12}\text{O}_6$	CH_2O
benzene	C_6H_6	CH
water	H_2O	H_2O

The molecular formula is always a whole-number multiple of the empirical formula.

■ Finding an empirical formula

1. Write the mass (or percentage) of each element.
2. Divide each by its A_r to get moles.
3. Divide every value by the smallest.
4. If a value ends in .5, .33 or .25, multiply all of them to get whole numbers.

element	% by mass	$\div A_r \rightarrow$ moles	$\div$ smallest (3.33)	ratio
C	40.0	$40.0/12 = 3.33$	1.0	1
H	6.7	$6.7/1 = 6.7$	2.0	2
O	53.3	$53.3/16 = 3.33$	1.0	1

empirical formula CH_2O - if $M_r = 180$, it fits $180/30 = 6$ times: molecular formula $\text{C}_6\text{H}_{12}\text{O}_6$

Divide by A_r , then by the smallest number of moles

■ Worked example: from masses to the molecular formula

A hydrocarbon contains 2.40 g of carbon and 0.60 g of hydrogen. Its M_r is 30.08.

- C: $2.40/12.01 = 0.200$ mol. H: $0.60/1.01 = 0.594$ mol.
- Divide by 0.200: C 1, H $2.97 \approx 3$. Empirical formula CH_3 , with $M_r = 15.04$.
- $30.08/15.04 = 2$, so the molecular formula is C_2H_6 .

■ Worked example: water of crystallization

2.50 g of blue hydrated copper(II) sulfate, $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$, is heated until it becomes white. 1.60 g of anhydrous CuSO_4 remains.

- Water lost: $2.50 - 1.60 = 0.90$ g; $0.90/18.02 = 0.0499$ mol.
- CuSO_4 : $1.60/159.62 = 0.0100$ mol.
- Ratio $0.0499/0.0100 = 4.99 \approx 5$, so $x = 5$: $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

Concentration of solutions

■ Molar concentration

A **solution** is a **solute** dissolved in a **solvent**. The **concentration** is the amount of solute per unit volume of solution.

$$c = \frac{n}{V}$$

c in mol dm^{-3} , n in mol, V in dm^3 . Remember $1 \text{ dm}^3 = 1000 \text{ cm}^3$, so divide a volume in cm^3 by 1000.

Concentration can also be given in g dm^{-3} : multiply mol dm^{-3} by the molar mass.

■ Worked example: making a solution

5.85 g of NaCl is dissolved in water and made up to 250 cm^3 .

- $n = 5.85/58.44 = 0.100$ mol.
- $V = 250/1000 = 0.250 \text{ dm}^3$.
- $c = 0.100/0.250 = 0.400 \text{ mol dm}^{-3}$, which is $0.400 \times 58.44 = 23.4 \text{ g dm}^{-3}$.

A **standard solution** has an accurately known concentration. Weigh the solid, dissolve it, transfer every drop to a volumetric flask, and make up to the mark with distilled water.

■ Dilution

Adding water does not change the amount of solute, only the volume:

$$c_1 V_1 = c_2 V_2$$

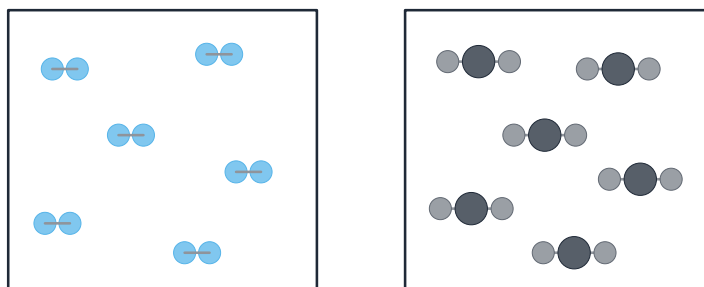
25.0 cm^3 of 2.00 mol dm^{-3} acid made up to 500 cm^3 gives $c_2 = 2.00 \times 25.0/500 = 0.100 \text{ mol dm}^{-3}$.

Avogadro's law

■ Equal volumes, equal numbers

Avogadro's law: equal volumes of all gases, at the same temperature and pressure, contain equal numbers of molecules.

same volume, same temperature and pressure



hydrogen, H₂

carbon dioxide, CO₂

the same **number** of molecules (6 each), even though CO₂ is bigger

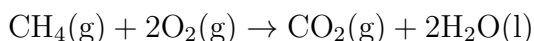
Gas molecules are far apart, so their size hardly affects the volume

It works because the molecules in a gas are very far apart. The volume is almost all empty space, so it depends on the number of molecules, not on their size.

At STP (273 K and 100 kPa) one mole of any ideal gas occupies 22.7 dm³ (data booklet).

■ Volumes of reacting gases

For gases at the same temperature and pressure, the volume ratio equals the mole ratio in the equation.



20 cm³ of methane needs 40 cm³ of oxygen and makes 20 cm³ of carbon dioxide. The water is a liquid at room temperature, so it is not counted as a gas volume.

Practice

1. Calculate the number of atoms in 0.20 mol of ammonia, NH₃.
2. Natural boron is 20% ¹⁰B and 80% ¹¹B. Calculate its *A_r*.
3. Calculate the molar mass of calcium hydroxide, Ca(OH)₂.
4. Calculate the amount in moles in 8.0 g of methane, CH₄.
5. Calculate the mass of 3.01×10^{23} molecules of oxygen, O₂.
6. Calculate the percentage by mass of iron in iron(III) oxide, Fe₂O₃.
7. A compound contains 85.6% carbon and 14.4% hydrogen by mass. Its *M_r* is 56.12. Find the empirical and molecular formulae.
8. 3.00 g of magnesium burns to give 4.97 g of oxide. Find the empirical formula of the oxide.

9. 4.80 g of hydrated magnesium sulfate, $\text{MgSO}_4 \cdot x\text{H}_2\text{O}$, loses 2.46 g of water on heating. Find x .
10. Calculate the concentration in mol dm^{-3} of 2.00 g of NaOH in 500 cm^3 of solution.
11. What volume of $0.500 \text{ mol dm}^{-3}$ HCl must be diluted to make 250 cm^3 of $0.0200 \text{ mol dm}^{-3}$ HCl?
12. 50 cm^3 of propane burns: $\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l})$. Calculate the volume of oxygen used and of carbon dioxide made, at the same temperature and pressure.

Answers.

1. Molecules = $0.20 \times 6.02 \times 10^{23} = 1.20 \times 10^{23}$; 4 atoms each, so 4.82×10^{23} atoms.
2. $(20 \times 10 + 80 \times 11)/100 = 10.8$.
3. $40.08 + 2 \times (16.00 + 1.01) = 74.10 \text{ g mol}^{-1}$.
4. $M = 16.05 \text{ g mol}^{-1}$; $n = 8.0/16.05 = 0.50 \text{ mol}$.
5. $n = 3.01 \times 10^{23}/6.02 \times 10^{23} = 0.500 \text{ mol}$; $m = 0.500 \times 32.00 = 16.0 \text{ g}$.
6. $M_r = 2 \times 55.85 + 3 \times 16.00 = 159.70$; % Fe = $111.70/159.70 \times 100 = 69.9\%$.
7. C: $85.6/12.01 = 7.13$; H: $14.4/1.01 = 14.3$; ratio 1 : 2, so CH_2 ($M_r = 14.03$). $56.12/14.03 = 4$, so C_4H_8 .
8. O: $4.97 - 3.00 = 1.97 \text{ g}$. Mg: $3.00/24.31 = 0.123 \text{ mol}$; O: $1.97/16.00 = 0.123 \text{ mol}$. Ratio 1 : 1, so MgO.
9. MgSO_4 : $4.80 - 2.46 = 2.34 \text{ g}$, $2.34/120.38 = 0.0194 \text{ mol}$. Water: $2.46/18.02 = 0.137 \text{ mol}$. $0.137/0.0194 = 7.0$, so $x = 7$.
10. $n = 2.00/40.00 = 0.0500 \text{ mol}$; $c = 0.0500/0.500 = 0.100 \text{ mol dm}^{-3}$.
11. $V_1 = 0.0200 \times 250/0.500 = 10.0 \text{ cm}^3$.
12. Oxygen $5 \times 50 = 250 \text{ cm}^3$; carbon dioxide $3 \times 50 = 150 \text{ cm}^3$.

Exam tips

- Say which particle you count: “0.5 mol of O_2 molecules” is 1 mol of O atoms.
- Use the data booklet A_r values and keep them to 2 decimal places; round only the final answer.
- Divide volumes in cm^3 by 1000 before using $c = n/V$.
- For an empirical formula, never round 1.5 or 1.33 to a whole number; multiply all the values instead.
- Gas volume ratios equal mole ratios only for gases at the same temperature and pressure; liquids and solids do not count.