

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.

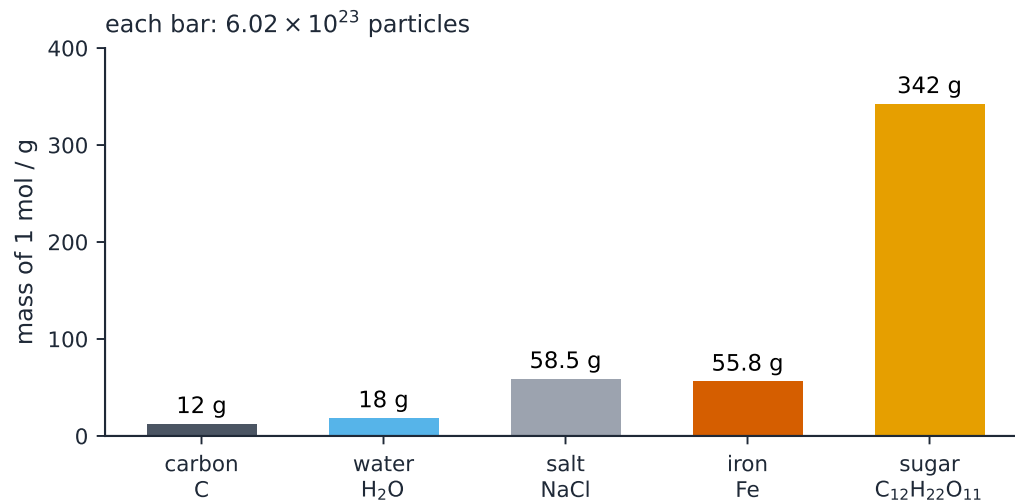
■ Relative formula mass

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} : $M(\text{CaCO}_3) = 100.09 \text{ g mol}^{-1}$.



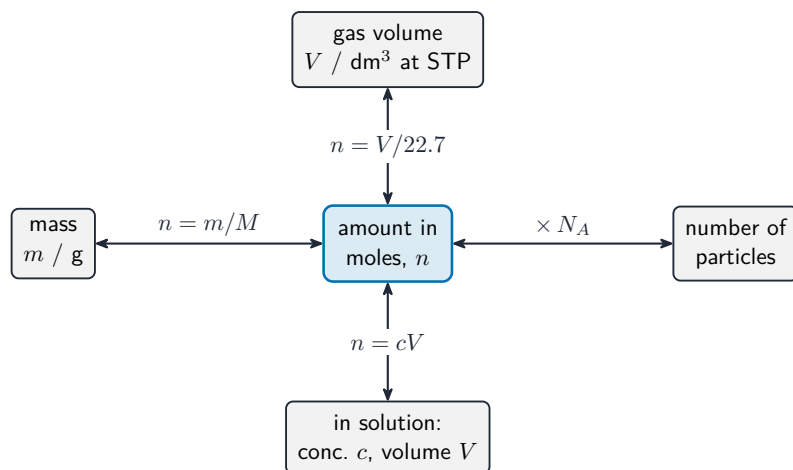
One mole of each substance has the same number of particles but a different mass

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?

- $M(\text{CO}_2) = 12.01 + 2 \times 16.00 = 44.01 \text{ g mol}^{-1}$.
- $n = 4.4/44.01 = 0.10 \text{ mol}$.
- $N = 0.10 \times 6.02 \times 10^{23} = 6.0 \times 10^{22}$ 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	$\div A_r \rightarrow \text{moles}$	$\div \text{smallest (3.33)}$	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: $2.40/12.01 = 0.200 \text{ mol}$. H: $0.60/1.01 = 0.594 \text{ mol}$.
- 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:

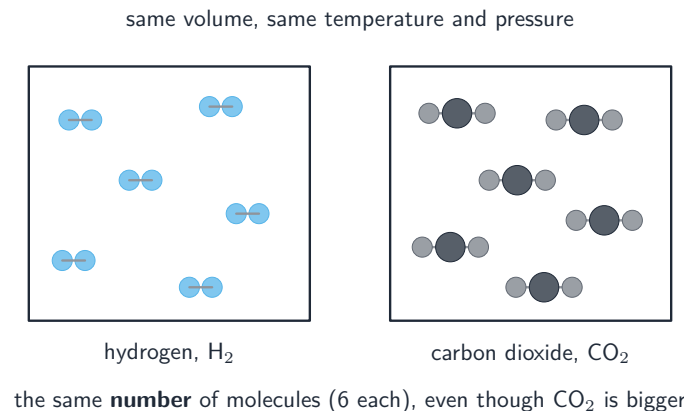
$$c_1V_1 = c_2V_2$$

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.



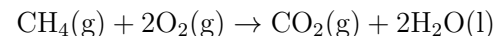
Gas molecules are far apart, so their size hardly affects the volume

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 dm^3 (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 cm^3 of methane needs 40 cm^3 of oxygen and makes 20 cm^3 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, NH_3 .
2. Natural boron is 20% ^{10}B and 80% ^{11}B . Calculate its A_r .
3. Calculate the molar mass of calcium hydroxide, $\text{Ca}(\text{OH})_2$.
4. Calculate the amount in moles in 8.0 g of methane, CH_4 .
5. Calculate the mass of 3.01×10^{23} molecules of oxygen, O_2 .
6. Calculate the percentage by mass of iron in iron(III) oxide, Fe_2O_3 .
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.