

Exercise walk-through

2.2

The mole and the Avogadro constant

A-Level Chemistry

Relative masses

Masses are compared with $\frac{1}{12}$ of a ^{12}C atom (the unified atomic mass unit, u).

 A_r

average mass
of an element's
atoms – weighted
over its isotopes

 M_r

sum of the A_r
values of every atom
in the molecule

formula mass

the same idea for
an ionic compound

Chlorine is 75% ^{35}Cl and 25% ^{37}Cl

$$A_r = (75 \times 35 + 25 \times 37)/100 = 35.5$$

Four relative masses, one reference atom

Relative isotopic mass

The mass of one atom of a *single* isotope, against $\frac{1}{12}$ of a carbon-12 atom.

Relative atomic mass

A_r : the weighted average over the isotopes as they occur – which is why it is rarely a whole number.

Relative molecular mass

M_r of a **molecule** : the sum of the relative atomic masses of all its atoms.

Relative formula mass

The same idea for a substance not made of molecules – an ionic compound, where there is no molecule to weigh.

The Avogadro constant

$6.02 \times 10^{23} \text{ mol}^{-1}$: how many particles are in a mole, and what turns any of these masses into a count.

All four are ratios against the same standard, so they carry **no units**. Only the Avogadro constant does.

The mole

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

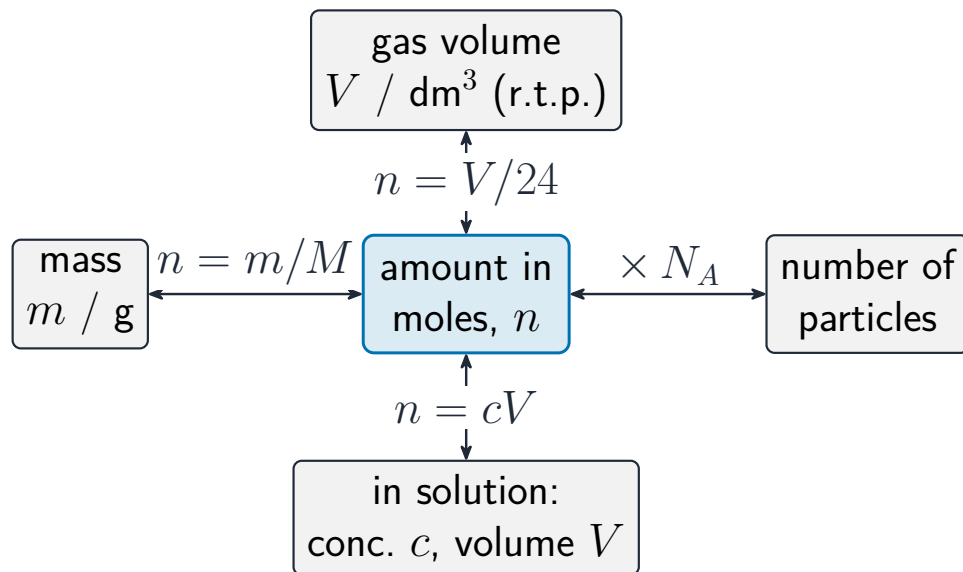
$$N_A = 6.02 \times 10^{23} \text{ mol}^{-1}$$

Worked example

moles in 8.0 g of CH_4 ($M = 16$): $n = \frac{8.0}{16} = 0.50 \text{ mol}$

One **mole** contains N_A particles (atoms, molecules *or* ions); the **molar mass** M in g mol^{-1} equals A_r or M_r .

The mole map



A **compound** is two or more elements chemically joined. Chemists count its particles in **moles**, as we count eggs in dozens – the hub of every amount calculation:

- mass: $n = m/M$
- particles: $\times N_A$
- gas: $n = V/24$
- solution: $n = cV$

Exercises 2.1·3.1

2.1 State the value of the Avogadro constant and what it counts.

[2]

2.2 Calculate the number of moles in 36 g of water, H_2O ($M = 18$).

[1]

3.1 How many molecules are in 2.0 mol of carbon dioxide?

[1]

☐

A 3.0×10^{23}

B 6.02×10^{23}

C 1.2×10^{24}

D 2.0×10^{23}

Exercises 2.1·3.1

2.1 State the value of the Avogadro constant and what it counts.

[2]

2.1

- $6.02 \times 10^{23} \text{ mol}^{-1}$; the number of particles (atoms/molecules/ions) in one mole

2.2 Calculate the number of moles in 36 g of water, H_2O ($M = 18$).

[1]

2.2

- $n = 36/18 = 2.0 \text{ mol}$

3.1 How many molecules are in 2.0 mol of carbon dioxide?

[1]

☐

A 3.0×10^{23}

B 6.02×10^{23}

C 1.2×10^{24}

D 2.0×10^{23}

3.1 C

- $2.0 \times 6.02 \times 10^{23} = 1.2 \times 10^{24}$

Exercises 3.2·3.3

3.2 The mass of 0.25 mol of a compound is 22 g. Its molar mass is: [1]

☐

A 5.5 g mol⁻¹

B 44 g mol⁻¹

C 88 g mol⁻¹

D 5.5×10^{23} g mol⁻¹

3.3 Calculate the number of moles in 5.85 g of sodium chloride, NaCl ($M = 58.5$). [2]

Exercises 3.2-3.3

3.2 The mass of 0.25 mol of a compound is 22 g. Its molar mass is: [1]

☐

A 5.5 g mol⁻¹

B 44 g mol⁻¹

C 88 g mol⁻¹

D 5.5×10^{23} g mol⁻¹

3.2 C

- $M = m/n = 22/0.25 = 88 \text{ g mol}^{-1}$

3.3 Calculate the number of moles in 5.85 g of sodium chloride, NaCl ($M = 58.5$). [2]

3.3

- $n = m/M = 5.85/58.5 = 0.100 \text{ mol}$

Exercise 3.4

3.4 Calculate the number of chloride ions in 5.85 g of NaCl ($N_A = 6.02 \times 10^{23}$). [2]

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3.4

- moles of NaCl = 0.100, so moles of Cl^- = 0.100; number = $0.100 \times 6.02 \times 10^{23} = 6.02 \times 10^{22}$