

Top tips

Quantitative chemistry

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This article gives practical tips for tackling chemistry calculations in exams

Exam links

- energetics
- equilibria

Numbers really matter in chemistry. Whether you are working out an equilibrium constant (K_c), finding an activation energy (E_a) or calculating an enthalpy change (ΔH), the maths behind the chemistry can make or break your answer. The good news? Most mistakes are not about how much chemistry you know – they are errors with rearranging equations, mishandling significant figures or forgetting a unit conversion. Here is a crash course to avoid those classic traps.

Rearranging equations

Rearranging is the foundation of quantitative chemistry: one wrong move and your whole calculation crumbles.

Example

The Arrhenius equation is:

$$k = A e^{-\frac{E_a}{RT}}$$

where E_a is the activation energy (J mol^{-1}), A is the pre-exponential (or Arrhenius) factor, R is the universal gas constant ($\text{J K}^{-1} \text{mol}^{-1}$) and T is the absolute temperature of the reaction (K).

You will often need to find the activation energy (E_a). Start by taking natural logs of both sides to 'reverse' the exponential that is present:

$$\ln k = \ln A - \left(\frac{E_a}{R}\right)\left(\frac{1}{T}\right).$$

This is now in the form $y = c + mx$, with gradient equal to $-\frac{E_a}{R}$.

This means that if you plot $\ln k$ (y-axis) against $\frac{1}{T}$ (x-axis) you can get E_a from the negative gradient (multiplying the gradient value by R , giving a positive E_a).

Exam tip

When rearranging, make sure the units cancel fully. Units will save your calculation.

For example, to determine the units of k in the Arrhenius equation:

$$k = A e^{-\frac{E_a}{RT}}$$

The units cancel out:

$$\frac{E_a}{RT} = \frac{\cancel{\text{mol}^{-1}}}{\cancel{\text{K}^{-1}} \cancel{\text{mol}^{-1}} \times \cancel{\text{K}}}$$

So, this part of the equation has no units:

$$k = A e^{-\frac{E_a}{RT}}$$

Consequently, units of k = units of A .

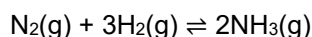
In equations, symbols in italics show variables or physical quantities (things we can measure). For example, A is often treated like a variable, because its value depends on the reaction. Symbols in upright type show defined constants, names or standard mathematical functions (e.g. e is a mathematical constant called Euler's number).

Standard form

Chemistry uses extreme numbers: tiny concentrations and giant constants. To make these numbers easier to handle, use standard form.

Example

For the Haber process:



$$K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$

with $[\text{N}_2] = 0.050 \text{ mol dm}^{-3}$, $[\text{H}_2] = 0.150 \text{ mol dm}^{-3}$, $[\text{NH}_3] = 0.0020 \text{ mol dm}^{-3}$.

Plugging in without standard form can lead to chaos. Use powers of 10 to keep it cleaner. For example:

$$K_c = \frac{[2.0 \times 10^{-3}]^2}{[5.0 \times 10^{-2}][1.50 \times 10^{-1}]^3}$$

For the units of the equilibrium constant K_c :

$$\frac{[\text{mol dm}^{-3}]^2}{\text{mol dm}^{-3} \times [\text{mol dm}^{-3}]^3}$$

which cancels to give $\text{mol}^{-2} \text{ dm}^6$.

Exam tip

Rewrite numbers in powers of 10 before calculating – it saves time and avoids confusion.

Significant figures

If you quote 6 significant figures when your data only gives 2, it is inaccurate and you will lose marks.

Example

Consider the question:

Calculate the heat energy, when 0.0250 mol of methane is burnt with an enthalpy change of $\Delta H = -890 \text{ kJ mol}^{-1}$.

The calculation is:

$$0.0250 \text{ mol} \times 890 \text{ kJ mol}^{-1} = 22.25 \text{ kJ}$$

But the limiting data has 3 significant figures, so the correct answer is 22.3 kJ (making sure the mol and mol^{-1} units cancel).

Exam tip

Match your significant figures to the least precise piece of data.

Units

Conversions trip up even strong students. Do not let units cost you marks. Common conversions include:

- $\text{cm}^3 \leftrightarrow \text{dm}^3$ ($1000 \text{ cm}^3 = 1 \text{ dm}^3$)
- $\text{J} \leftrightarrow \text{kJ}$ ($1000 \text{ J} = 1 \text{ kJ}$).

Example

In calorimetry, 5200 J are released when 0.200 mol of a compound reacts. The enthalpy change equals:

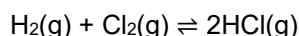
$$\frac{5200 \text{ (J)}}{0.200 \text{ (mol)}}$$

giving ΔH as $-26000 \text{ J mol}^{-1}$. However, this enthalpy change should be converted to $-26.0 \text{ kJ mol}^{-1}$.

Exam tip

Before you start, circle all the units in the question. Check them and make sure they cancel out before you write your final answer (including final units). For example:

Calculate K_c of the following reaction



The equilibrium concentrations are:

$$[\text{H}_2] = 0.67 \text{ mol dm}^{-3} \quad [\text{Cl}_2] = 0.83 \text{ mol dm}^{-3} \quad \text{and} \quad [\text{HCl}] = 0.33 \text{ mol dm}^{-3}$$

Putting it all together

When tackling chemistry calculations, always work step by step: set up the equation, substitute the values, check the units and only then calculate. A stepwise approach reduces errors and makes it easy to pick up method marks even if the final answer is wrong.

Example

Consider the question:

For $2\text{NO}_2 \rightleftharpoons \text{N}_2\text{O}_4$ at 298 K, $[\text{NO}_2] = 0.020 \text{ mol dm}^{-3}$ and $[\text{N}_2\text{O}_4] = 0.0050 \text{ mol dm}^{-3}$.

Calculate K_c .

- Step 1: give the equation, $K_c = \frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2}$

- Step 2: substitute in the values, $\frac{[5.0 \times 10^{-3}]}{[2.0 \times 10^{-2}]^2} = 12.5 \text{ mol}^{-1} \text{ dm}^3$.
In terms of units, $\frac{[\text{mol dm}^{-3}]}{[\text{mol dm}^{-3}]^2} = \frac{[\text{mol dm}^{-3}]}{[\text{mol}^2 \text{ dm}^{-6}]} = \text{mol}^{-1} \text{ dm}^3$.
- Step 3: finally, round to the correct number of significant figures, which is $13 \text{ mol}^{-1} \text{ dm}^3$.

Exam tip

Always check the number of significant figures last. Do all your calculations first, then round at the very end.

Here is a checklist to follow:

- Rearrange the equation clearly, step by step.
- Convert numbers to standard form (if needed) before substituting.
- Identify which value has the fewest significant figures.
- Convert all units to the correct form (e.g. $\text{cm}^3 \rightarrow \text{dm}^3$, $\text{J} \rightarrow \text{kJ}$).
- Check that units cancel correctly to give the expected final units.
- Round your answer only at the end.

Final word

Mastering quantitative chemistry is less about being a maths genius and more about being methodical. Rearrange carefully, use standard form, match your significant figures and check your units. Do that, and you will not only protect your marks – you will make the chemistry itself feel much clearer. Numbers do not have to be the enemy. With a few smart habits, they will become your secret weapon.

Practice exam questions

- 1 1 kg of water is cooled by 15.0°C . Given that the specific heat capacity is $4.18 \text{ J g}^{-1} \text{ K}^{-1}$, calculate the energy released. Give your answer in kJ in standard form. (2 marks)
- 2 Calculate the number of moles in 2.45 grams of sodium carbonate, Na_2CO_3 . (1 marks)
- 3 Find the mass of 200 cm^3 of carbon monoxide (CO) at a temperature of 34°C and a pressure of 97.9 kPa. The gas constant is $8.31 \text{ J mol}^{-1} \text{ K}^{-1}$. (2 marks)
- 4 Given Avogadro's constant is $6.02 \times 10^{23} \text{ mol}^{-1}$, calculate the number of hydrogen atoms in 1.00 g of water. (2 marks)
- 5 Find the volume of 2.0 g of O_2 when it is heated to 60°C at 1 atm of pressure. Give your answer in dm^3 , to 2 significant figures. (1 atm = 101325 Pa and the gas constant is $8.31 \text{ J mol}^{-1} \text{ K}^{-1}$.) (2 marks)

Answers are on the next page

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Answers

1 $q = mc\Delta T$

$$= 1000 \times 4.18 \times 15.0 = 62700 \text{ J}$$

In terms of units: $1000 \text{ (g)} \times 4.18 \text{ (J g}^{-1} \text{ K}^{-1}) \times 15.0 \text{ (K)} = 62700 \text{ J}$

As requested in the question, convert to kJ and standard form: $62700 \text{ J} = 6.27 \times 10^1 \text{ kJ}$

2 $M_r = (23 \times 2) + 12 + (16 \times 3) = 106$ (no units as it is a relative mass)

Molar mass of $\text{Na}_2\text{CO}_3 = 106 \text{ g mol}^{-1}$

Moles = mass / molar mass (M_r)

$$= 2.45 \text{ g} / 106 \text{ g mol}^{-1} = 0.023 \text{ mol}$$

3 Temperature in Kelvin = $34 + 273 = 333 \text{ K}$

Rearrange $n = PV / RT$

$P = 97900 \text{ Pa}$; $V = 2 \times 10^{-4} \text{ m}^3$; $T = 307 \text{ K}$

$$n = (97900 \text{ Pa} \times 2 \times 10^{-4} \text{ m}^3) / (8.31 \text{ J mol}^{-1} \text{ K}^{-1} \times 307 \text{ K}) = 7.674... \times 10^{-3} \text{ moles}$$

In terms of units:

$$\frac{\text{Pa m}^3}{\text{J mol}^{-1}} = \frac{\text{kg m}^{-1} \text{ s}^{-2} \text{ m}^3}{\text{kg m}^2 \text{ s}^{-2} \text{ mol}^{-1}} = \frac{\text{kg m}^2 \text{ s}^{-2}}{\text{kg m}^2 \text{ s}^{-2} \text{ mol}^{-1}} = \frac{1}{\text{mol}^{-1}} = \text{mol}$$

$$\text{mass} = 7.674... \times 10^{-3} \times 28 \text{ g mol}^{-1} = 0.21 \text{ g}$$

4 Moles of water molecules = $1.00 \text{ g} / 18 \text{ g mol}^{-1} = 0.0556 \text{ mol}$

Number of water molecules = $0.0556 \text{ mol} \times 6.02 \times 10^{23} \text{ mol}^{-1} = 3.34 \times 10^{22}$ (no units)

Each water molecule has two hydrogens, so the number of hydrogen atoms:

$$3.34 \times 10^{22} \times 2 = 6.68 \times 10^{22}$$

5 Number of moles of O_2 = mass / molar mass = $2.0 \text{ g} / 32 \text{ g mol}^{-1} = 0.0625 \text{ mol}$

Temperature in Kelvin = $60 + 273 = 333 \text{ K}$.

Rearranging $n = PV / RT$, gives $V = nRT / P$

$$V = (0.0625 \text{ mol} \times 8.31 \text{ J mol}^{-1} \text{ K}^{-1} \times 333 \text{ K}) / 101325 \text{ Pa}$$

$$= 1.71 \times 10^{-3} \text{ m}^3 = 1.71 \text{ dm}^3$$

In terms of units:

$$\frac{\text{mol J mol}^{-1} \text{ K}^{-1} \text{ K}}{\text{Pa}} = \frac{\text{mol kg m}^2 \text{ s}^{-2} \text{ mol}^{-1} \text{ K}^{-1} \text{ K}}{\text{kg m}^{-1} \text{ s}^{-2}} = \text{m}^3$$