The Chemical Industry (CI)

Revision guide

Clarity, confidence, and top grades.

What this storyline covers

- Rate equations, orders and the rate constant
- Finding orders from experiments
- Half-life and the rate-determining step
- The Arrhenius equation
- Equilibrium and K_c calculations (with units)
- Nitrogen chemistry
- Industrial conditions, costs and risks

Required practicals in this storyline: reaction kinetics (**PAG9/10**), measuring K_c, and tests for nitrate(V) and ammonium ions.



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i How to use this guide

The Chemical Industry is the **quantitative heart** of the A2 course. It develops **reaction kinetics** (rate equations, orders, the rate constant and the Arrhenius equation) and **equilibrium** (K_c calculations, now with units), then applies both to choosing the best conditions for an industrial process. Nitrogen chemistry provides the context, and the kinetics and K_c work carries straight into Oceans.

How it is assessed

OCR B is context-led, so CI content can appear in **any** of the three written papers, in any context:

- **Paper 1** — Fundamentals of chemistry (110 marks)
- **Paper 2** — Scientific literacy, including the pre-release Advance Notice article (100 marks)
- **Paper 3** — Practical skills in chemistry (60 marks)

How to revise it

- Learn the **definitions in the teal boxes** word-for-word — they are marks you can guarantee.
- Work every **worked example** yourself — this storyline is calculation-heavy.
- Practise **deducing orders and units** until they are automatic.
- Use the **spec checklist** on the last page to confirm you can do every point.

1 Rate equations and orders

Spec reference: CI(a).

Key definitions

- **Rate of reaction** — the change in concentration of a reactant or product per unit time (mol dm^{-3} s^{-1}).
- **Order of reaction** (with respect to a reactant) — the power to which that reactant's concentration is raised in the rate equation. The **overall order** is the sum of the individual orders.
- **Rate constant, k** — the constant in the rate equation; it is found experimentally and **increases with temperature**.

$$\text{rate} = k[A]^m[B]^n$$

m and n are the orders (integers 0, 1 or 2); the units of k depend on the overall order

- '$\propto$' means 'proportional to': if a reaction is first order in A, then rate $\propto$ [A].
- Find the **units of k** by rearranging to k = rate ÷ (concentration terms), then cancelling.

From these initial-rate data, deduce the rate equation:

Experiment	[A] / mol dm^{-3}	[B] / mol dm^{-3}	rate / mol dm^{-3} s^{-1}
1	0.10	0.10	0.20
2	0.20	0.10	0.40
3	0.10	0.20	0.80

Worked example — order, rate constant and units

1 → 2: [A] × 2, rate × 2 → first order in A
1 → 3: [B] × 2, rate × 4 → second order in B
rate = $k[A][B]^2$ (overall third order)
$k = 0.20 / (0.10 \times 0.10^2) = 200$ mol^{-2} dm^6 s^{-1}

2 Finding orders from experiments

Spec references: CI(b), CI(c). Practical: reaction kinetics (PAG9/10).

Two approaches CI(c)

- **Initial rates:** measure the initial rate for different starting concentrations; see how the rate changes when **one** concentration is varied to find the order in that reactant.
- **Continuous monitoring:** follow a concentration against time and read rates from the **tangents** to the curve, or use **half-lives**.

Recognising the order from a graph CI(c)

Order	Concentration–time graph	Rate–concentration graph
Zero	straight line (constant gradient)	horizontal line
First	curve with a **constant half-life**	straight line through the origin
Second	steeper curve; the half-life increases	upward curve

Half-life CI(b)

- The **half-life** ($t_{½}$) is the time for a reactant's concentration to halve.
- A **constant** half-life is the signature of a **first-order** reaction; for first order, $t_{½} = \ln 2 / k = 0.693 / k$.

Worked example — half-life

A first-order reaction has $k = 0.0231\ s^{-1}$. Find its half-life.

$t_{½} = 0.693 / k = 0.693 / 0.0231 = 30.0\ s$

3 The rate-determining step

Spec reference: CI(e).

Key definition

The **rate-determining step** is the **slowest step** in a multi-step mechanism; it controls the overall rate.

- Only the species involved **up to and including** the rate-determining step appear in the rate equation.
- The **order** with respect to a reactant tells you how many of its molecules are involved up to that step, so the orders provide **evidence for the mechanism**.
- Example: if a reaction is **first order in A** and **zero order in B**, then one molecule of A — and no B — is involved in the rate-determining step; B must react in a later, faster step.

4 The Arrhenius equation

Spec reference: CI(d).

$$k = Ae^{-E_a/RT} \quad \text{or} \quad \ln k = \ln A - E_a/RT$$

A = pre-exponential factor; E_a = activation energy; $R = 8.314\ J\ mol^{-1}\ K^{-1}$; T = temperature (K). The equation is given on the Data Sheet.

- The equation shows that k (and so the rate) **increases with temperature** and is **smaller** when E_a is larger.
- To find E_a and A, plot **ln k against 1/T**: the line has **gradient = $-E_a/R$** and **intercept = ln A**.

Worked example — activation energy from a gradient

A plot of ln k against 1/T has a gradient of -6.0×10^3 K. Find E_a.

gradient = $-E_a/R$, so E_a = −gradient × R

$E_a = 6.0 \times 10^3 \times 8.314 = 4.99 \times 10^4\ J\ mol^{-1} = 49.9\ kJ\ mol^{-1}$

5 Equilibrium and K_c

Spec references: CI(f), CI(g), CI(h). Practical: measuring K_c.

$$\text{for } aA + bB \rightleftharpoons cC + dD, \quad K_c = [C]^c[D]^d \div [A]^a[B]^b$$

equilibrium concentrations in mol dm^{-3}; here the units must be worked out

- **Units of K_c:** substitute mol dm^{-3} for each concentration and cancel. They may cancel completely (no units) or leave a power of mol dm^{-3}.

Worked example — K_c with units

For $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ at equilibrium: $[N_2] = 0.20$, $[H_2] = 0.20$, $[NH_3] = 0.40$ mol dm^{-3}.

$K_c = [NH_3]^2 / ([N_2][H_2]^3) = (0.40)^2 / (0.20 \times 0.20^3)$

$= 0.16 / (0.20 \times 0.0080) = 0.16 / 0.0016 = 100$

units = $(mol\ dm^{-3})^2 \div (mol\ dm^{-3})^4 = mol^{-2}\ dm^6$, so $K_c = 100\ mol^{-2}\ dm^6$

What changes K_c CI(f)

- **Temperature** changes the value of K_c (for an exothermic reaction, raising T **decreases** K_c; for an endothermic one it increases it).
- **Pressure** does **not** change K_c; it only shifts the position for gas reactions.
- A **catalyst** has **no effect** on K_c or the position of equilibrium — it only speeds up how fast equilibrium is reached.

The best conditions for industry CI(g)

The most economical conditions **balance yield against rate and cost**. For the **Haber process**, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ (exothermic):

- **High pressure** increases the yield (fewer gas molecules on the right) but the plant is expensive, so a compromise of about **200 atm** is used.
- **Low temperature** increases the yield (the reaction is exothermic) but is too slow, so a compromise of about **450 °C** with an **iron catalyst** gives an acceptable rate.

6 Nitrogen chemistry

Spec reference: CI(j).

Bonding CI(j)

- **Nitrogen gas, N_2:** a **triple bond** N≡N (one σ and two π bonds); very strong, so N_2 is unreactive.
- **Ammonia, NH_3:** three N–H bonds and one **lone pair**; pyramidal.
- **Ammonium ion, NH_4^+:** the lone pair on NH_3 forms a **dative (coordinate) bond** to H^+, giving four identical N–H bonds (tetrahedral, charge +1).

Oxides of nitrogen and interconversions CI(j)

- The oxides are **N_2O** (dinitrogen oxide, colourless), **NO** (nitrogen monoxide, colourless) and **NO_2** (nitrogen dioxide, brown).
- Nitrogen takes a range of oxidation states, so the **ammonium ion (−3)**, the oxides, the **nitrate(III) ion NO_2^- (+3)** and the **nitrate(V) ion NO_3^- (+5)** interconvert by oxidation and reduction.

Tests for nitrate(V) and ammonium ions CI(j)

- **Nitrate(V), NO_3^-:** warm with **Devarda's alloy and sodium hydroxide** — ammonia gas is evolved (it turns damp red litmus blue).
- **Ammonium, NH_4^+:** warm with **sodium hydroxide** — ammonia gas is evolved (it turns damp red litmus blue).

7 Industrial processes: chemistry, costs and risks

Spec references: CI(i), CI(k).

The reactions CI(i)

- You are expected to **interpret given equations** for industrial processes; no recall is required. The storyline uses the manufacture of **nitric acid**, **sulfuric acid** and **ethanoic acid**.

Choosing conditions and weighing costs CI(g), CI(k)

- The **best operating conditions** balance **equilibrium yield**, **reaction rate** and **cost** (see §5).
- **Costs** to weigh include **raw materials**, **energy**, the **plant** itself, and the value of any **co-products and by-products**.
- **Benefits and risks:** weigh the **benefit to society** of the product against the **hazards** of the process (toxicity, explosion risk, environmental impact). You analyse the information given.

P Required practicals in this storyline

These PAG activities are most likely to be examined in Paper 3, but the techniques can appear in any paper.

Activity	What to be able to do
Reaction kinetics (PAG9/10)	follow a reaction over time; use initial rates or graphical methods (tangents, half-lives) to find the order, the rate constant and the activation enthalpy, and construct a rate equation
Measuring K_c	carry out an equilibrium reaction, measure the equilibrium concentrations, and calculate K_c with its units
Tests for nitrogen compounds	warm a nitrate(V) with Devarda's alloy and NaOH, and an ammonium salt with NaOH; in each case identify the ammonia evolved

Exam tip — the calculation marks

Deduce orders by comparing experiments where **only one** concentration changes. Always work out the **units of k and K_c** by substituting mol dm^{-3} and cancelling. When asked for industrial conditions, **state the trade-off** — yield against rate against cost.

Spec coverage checklist — can you do every point?

Every assessable learning outcome for The Chemical Industry, with where it is covered in this guide. Tick each one once you can do it confidently.

CI(a)	Rate of reaction, rate constant (with units), order, and the rate equation rate = $k[A]^m[B]^n$ — §1
CI(b)	Calculating half-lives from given data — §2
CI(c)	Finding order, rate constant and rate equation from experimental data (graphs, half-lives, initial rates) — §2
CI(d)	The Arrhenius equation; finding E_a and A — §4
CI(e)	The rate-determining step and its relation to orders and mechanism — §3
CI(f)	Effect of temperature, pressure and catalysts on K_c — §5
CI(g)	Choosing the most economical conditions for an industrial process — §5, §7
CI(h)	K_c calculations with units for homogeneous equilibria — §5
CI(i)	Interpreting the reactions in industrial processes — §7
CI(j)	Nitrogen chemistry: bonding, oxides, interconversions and tests — §6
CI(k)	Costs, benefits and risks of an industrial process — §7