Reactivity 2: How Much, How Fast and How Far?

IB SL Study Guide

Contents

💡 IB TIP

Data booklet: You can use the IB Chemistry Data Booklet in the exam — all constants, the periodic table, and key equations are provided.

IB Chemistry SL — Rates of Reaction & Equilibrium

Complete Study Guide

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1. Collision Theory & Activation Energy

Reactivity 2.2 — How fast? Rates of Reaction

The **rate of reaction** is a measure of how quickly reactants are converted into products. More precisely:

$$\text{Rate of reaction} = \frac{\Delta[\text{concentration}]}{\Delta t}$$

The units of rate are typically $\text{mol dm}^{-3}\text{ s}^{-1}$ (moles per decimetre cubed per second). The concentration of reactants decreases over time (so the rate for reactants is technically negative, but we report it as a positive number), while the concentration of products increases.

Collision Theory

For a chemical reaction to occur, particles must **collide**. But not every collision leads to a reaction — only **effective collisions** produce products. An effective collision requires two conditions to be met simultaneously:

1. **Sufficient energy** — the colliding particles must together have energy equal to or greater than the **activation energy** (E_a)
2. **Correct orientation** — the particles must collide at the right angle so that the reactive parts of the molecules interact

MEMORISE THIS

Collision theory in one sentence: Rate of reaction depends on the frequency of effective collisions — those with sufficient energy AND correct orientation.

- More collisions per second → more chance of effective collisions → faster rate
- Higher energy collisions → more particles exceed E_a → faster rate
- Better orientation → more productive collisions → faster rate

Activation Energy and the Maxwell-Boltzmann Distribution

The **activation energy** (E_a) is the minimum energy that colliding particles must possess for a reaction to occur. Think of it as an energy barrier that reactants must climb before they can become products.

Not all molecules in a sample have the same energy at a given temperature. The **Maxwell-Boltzmann distribution** shows how the energies of molecules are spread across a population at a fixed temperature.

Key features of the Maxwell-Boltzmann distribution curve:

Feature	Description
x-axis	Kinetic energy of molecules
y-axis	Number of molecules (or fraction of molecules) with that energy
Shape	Starts at zero (no molecules have zero energy), rises to a peak, then tails off to the right — the tail never reaches zero
Peak	Represents the most probable energy
Mean energy	Slightly to the right of the peak
E_a line	A vertical line on the x-axis; only molecules to the RIGHT of this line have enough energy to react
Shaded area	The area under the curve to the right of E_a represents the proportion of molecules that CAN react

What happens when temperature increases:

When temperature rises, the distribution curve changes in two important ways:

- The peak shifts to the **right** (higher energy) and flattens slightly (fewer molecules at the most probable energy)
- The tail extends further to the right
- **The area under the curve to the right of E_a increases significantly** — meaning a much larger proportion of molecules now have sufficient energy to react

This is the fundamental reason why increasing temperature increases the rate: more molecules exceed E_a, so more collisions are effective.

⚠ **EXAM ALERT**

A very common IB question asks you to explain, using the Maxwell-Boltzmann distribution, why a small increase in temperature causes a large increase in rate. The key phrase is: “a greater proportion of molecules now have energy greater than or equal to the activation energy.” Do not just say “molecules move faster” — you must link it to E_a.

IB TIP

You will not be asked to draw the Maxwell-Boltzmann distribution from scratch in a paper 1 MCQ, but you may be asked to interpret a pre-drawn diagram. Know: (1) what the x-axis and y-axis represent, (2) where E_a sits, (3) how the curve changes with temperature, and (4) how a catalyst shifts E_a (see Section 3).

WORKED EXAMPLE

Worked Example 1.1 — Reading the Maxwell-Boltzmann Distribution

A reaction has an activation energy of 50 kJ/mol. At 25 °C, 8% of molecules have energy $\geq E_a$. At 35 °C, 16% of molecules have energy $\geq E_a$.

Q: By what factor has the rate approximately increased?

Solution:

The fraction of molecules exceeding E_a has doubled (from 8% to 16%). Since all other factors (concentration, surface area) are unchanged, the rate of effective collisions approximately doubles.

Rate factor ≈ 2

This illustrates the general rule of thumb: a 10 °C rise roughly doubles the rate for many reactions — though this depends on E_a and is not a precise universal law.

2. Factors Affecting Rate of Reaction

Five key factors affect the rate of a chemical reaction. For each, the explanation must be in terms of **collision theory** — specifically, how the factor changes the **frequency** or **energy** of effective collisions.

Summary Table

Factor	Change	Effect on Rate	Collision Theory Explanation
Temperature	Increase	Increases	More molecules have energy ≥ E_a; collisions are more frequent AND more energetic
Concentration (solutions)	Increase	Increases	More particles per unit volume → more frequent collisions
Pressure (gases)	Increase	Increases	Equivalent to increasing concentration of gas particles
Surface area (solids)	Increase (smaller pieces)	Increases	More particles exposed at the surface → more collisions possible
Catalyst	Add catalyst	Increases	Provides alternative pathway with lower E_a (see Section 3)

Temperature

Increasing temperature gives molecules more kinetic energy. This has two effects:

1. Molecules move faster, so collisions occur more frequently
2. More importantly, a larger proportion of molecules now exceed E_a (seen as increased area under Maxwell-Boltzmann curve to the right of E_a)

Both effects increase the rate, but the increase in the fraction of molecules exceeding E_a is the dominant explanation in IB exams.

Concentration

In a solution, increasing the concentration means more solute particles are present in the same volume of solvent. Particles are closer together and collide more frequently. Since the frequency of collisions increases, the frequency of effective collisions also increases, and the rate rises.

Pressure (for gases)

Increasing pressure on a gas squeezes the same number of gas molecules into a smaller volume — this is equivalent to increasing concentration. Particles are closer together, collide more frequently, and the rate increases.

> **IB TIP**
> Pressure only matters for reactions involving gases. For reactions in solution, use “concentration” not “pressure.”

Surface Area

A solid reactant can only react at its surface — interior particles are not exposed. Breaking a solid into smaller pieces (e.g. powder vs lump) dramatically increases the total surface area without changing the amount of substance. More surface particles are exposed to collisions with the other reactant, increasing collision frequency and therefore rate.

EXAM ALERT

A classic IB application is explaining why fine flour dust in a mill can cause explosions (enormous surface area → extremely fast combustion rate), or why iron filings react with acid much faster than an iron nail.

WORKED EXAMPLE

Worked Example 2.1 — Calculating Average Rate

The concentration of a reactant A was measured over time:

Time / s	$[\mathrm{A}]$ / mol dm^{-3}
0	0.80
20	0.60
40	0.44
60	0.32
80	0.24

Q(a): Calculate the average rate of reaction between $t = 0$ and $t = 40$ s.

Q(b): Calculate the average rate between $t = 40$ s and $t = 80$ s.

Q(c): What does the change in rate tell you about the reaction as it proceeds?

Solution (a):

$$\text{average rate} = \frac{\Delta[\mathrm{A}]}{\Delta t} = \frac{0.80-0.44}{40-0} = \frac{0.36}{40} = 9.0 \times 10^{-3}\ \mathrm{mol\ dm^{-3}\ s^{-1}}$$

Solution (b):

$$\text{average rate} = \frac{0.44-0.24}{80-40} = \frac{0.20}{40} = 5.0 \times 10^{-3}\ \mathrm{mol\ dm^{-3}\ s^{-1}}$$

Solution (c): The rate has decreased (from 9.0×10^{-3} to 5.0×10^{-3} mol dm^{-3} s^{-1}). This is because as the reaction proceeds, the concentration of A decreases. Fewer particles per unit volume means fewer collisions per second, so the rate falls. This is why concentration-time graphs are curves (decreasing gradient) rather than straight lines.

3. Catalysts

Definition and Action

A **catalyst** is a substance that increases the rate of a chemical reaction without being consumed in the overall process. The catalyst is regenerated at the end of the reaction — it participates in the mechanism but is not used up.

Catalysts increase rate by providing an **alternative reaction pathway** that has a **lower activation energy** (E_a). This is the key IB explanation. The catalyst does NOT:

- Change the energy of the reactants
- Change the energy of the products
- Change the enthalpy change of reaction (ΔH)
- Shift the position of equilibrium (it speeds up forward and reverse equally)

On a Maxwell-Boltzmann diagram, adding a catalyst shifts the E_a line to the **left** (lower energy). This means a larger proportion of molecules now have energy ≥ E_a — the shaded area under the curve to the right of the new E_a is larger — so more collisions are effective and the rate increases.

MEMORISE THIS

How to explain a catalyst on the Maxwell-Boltzmann diagram:

"The catalyst provides an alternative reaction pathway with a lower activation energy. On the Maxwell-Boltzmann distribution, the E_a threshold shifts to the left. A greater proportion of molecules now have energy greater than or equal to this lower E_a, so a greater fraction of collisions are effective, and the rate increases."

Homogeneous vs Heterogeneous Catalysts

Type	Definition	Example
Homogeneous	Catalyst and reactants are in the **same phase** (same physical state)	$H^+(aq)$ catalysing ester hydrolysis in aqueous solution; $NO(g)$ catalysing oxidation of $SO_2(g)$ in the atmosphere
Heterogeneous	Catalyst and reactants are in **different phases**	Iron (solid) catalyst in the Haber Process; Pt (solid) in catalytic converter; V_2O_5 (solid) in Contact Process

Heterogeneous catalysts work by **adsorption** — reactant molecules are adsorbed onto the catalyst surface, weakening bonds and reducing E_a. After reaction, products desorb from the surface, freeing active sites for more reactant molecules.

Enzyme Catalysts

Enzymes are biological catalysts — large protein molecules that catalyse specific biochemical reactions in living organisms. Each enzyme has an **active site** with a specific shape that fits only one (or a few) substrate molecules (the **lock-and-key model**). Enzymes are extremely efficient and specific, operating under mild conditions (body temperature ~37 °C, near-neutral pH). In the IB, you should know that enzymes are:

- Biological (protein-based)
- Homogeneous catalysts (both enzyme and substrate are typically in aqueous solution)
- Highly specific (one enzyme, one substrate)
- Sensitive to temperature and pH (denaturation above ~40 °C)

⚠ **EXAM ALERT**

Do not confuse the effect of a catalyst on K_c with its effect on rate. A catalyst increases rate but does **not** change K_c or the equilibrium position — it helps the system reach equilibrium faster, from either direction.

WORKED EXAMPLE

Worked Example 3.1 — Catalyst Explanation

Q: A reaction A $\rightarrow$ B has $E_a = 120$ kJ mol^{-1}. A catalyst reduces this to 80 kJ mol^{-1}.

Explain, with reference to the Maxwell-Boltzmann distribution, why the rate increases.

Answer:

The catalyst provides an alternative reaction pathway with a lower activation energy (80 kJ mol^{-1} instead of 120 kJ mol^{-1}). On the Maxwell-Boltzmann energy distribution diagram, the activation energy threshold shifts to the left. A significantly greater proportion of molecules now have kinetic energy greater than or equal to 80 kJ mol^{-1} compared to 120 kJ mol^{-1}. This means a greater fraction of collisions are effective, so the rate of reaction increases. The overall energy of the reactants and products, and therefore ΔH, is unchanged.

4. Measuring Rate of Reaction

Rate Expression

Rate of reaction can be defined as:

$$\text{Rate} = \frac{\Delta[\text{concentration}]}{\Delta t}$$

In practice, “concentration of a species” can be substituted by any quantity proportional to concentration:

- **Volume of gas produced** (collected over water in a gas syringe or inverted measuring cylinder)
- **Mass lost** (if a gas is produced and escapes — measure on a balance)
- **Colour intensity** (if a coloured species is produced or consumed — use a colorimeter)
- **Conductivity** (if ions are produced or consumed)
- **Optical rotation** (for certain reactions involving chiral molecules)

Initial Rate vs Average Rate

Concept	Definition	How to find from graph
Average rate	Total change in concentration divided by total time elapsed	Gradient of the straight line connecting two points on a concentration-time curve
Initial rate	Rate at time $t = 0$ (when the reaction just starts)	Gradient of the tangent drawn at $t = 0$ on a concentration-time curve

The **initial rate** is the fastest rate during the reaction (for reactions where concentration of reactants decreases over time), because the reactant concentrations are at their highest at $t = 0$.

Rate-Concentration Graphs (Qualitative)

While IB SL does **not** require knowledge of rate laws or rate orders mathematically, you should be able to interpret graphs qualitatively:

Graph Shape	What it means
Rate stays constant as concentration changes (horizontal line)	Zero order — rate is independent of this reactant's concentration
Rate increases linearly with concentration (straight line through origin)	First order — rate is directly proportional to concentration
Rate increases as a curve (exponential-like) with concentration	Second order — rate is proportional to concentration squared

> **IB TIP**
>
> At SL, you will not be asked to write rate equations like “rate = k[A][B]” or to determine orders mathematically. You just need to know: (1) how to calculate average rate from data, (2) that rate decreases as a reaction proceeds (decreasing reactant concentration), and (3) how to interpret simple rate graphs qualitatively.

WORKED EXAMPLE

Worked Example 4.1 — Reading Rate from a Graph

A reaction produces gas. The volume of gas collected is measured over time:

Time / s	Volume of gas / cm^3
0	0
10	18
20	32
30	42
40	48
50	51
60	52

Q(a): Calculate the average rate of gas production between $t = 0$ and $t = 20$ s.

Q(b): Between $t = 40$ s and $t = 60$ s, the rate is much slower. Suggest why.

Q(c): At what point does the reaction appear to stop? Explain.

Solution (a):

$$\text{average rate} = \frac{32-0}{20-0} = \frac{32}{20} = 1.6\ \text{cm}^3\ \text{s}^{-1}$$

Solution (b): Between $t = 40$ and $t = 60$ s, only 4 cm^3 more gas is produced (compared to 32 cm^3 in the first 20 s). By this point, the reactants have been largely consumed — much lower concentration means fewer collisions per second, so the rate has fallen dramatically.

Solution (c): The reaction appears to stop at approximately $t = 55$–60 s, where the volume becomes constant at ~52 cm^3. This means no more gas is being produced — either a reactant has been fully consumed (limiting reagent exhausted) or, if this is a reversible reaction, equilibrium has been reached.

5. Dynamic Equilibrium & Le Chatelier's Principle

Reactivity 2.3 — How Far? Equilibrium

Many reactions are **reversible** — products can react to re-form reactants. We write reversible reactions with a double arrow ($\rightleftharpoons$):

$$\text{A} + \text{B} \rightleftharpoons \text{C} + \text{D}$$

What Is Dynamic Equilibrium?

At the start of a reversible reaction, only reactants are present. As they react, products build up. The **forward reaction** slows (as reactant concentrations fall) and the **reverse**

reaction speeds up (as product concentrations rise). Eventually, a state is reached where:

- The **rate of the forward reaction equals the rate of the reverse reaction**
- The **concentrations of all species remain constant** (but are NOT necessarily equal)
- The reaction is still occurring in both directions — it is **dynamic** (not static)

This state is called **dynamic equilibrium**.

MEMORISE THIS

Dynamic equilibrium — three things that are true:

1. Forward rate = reverse rate
2. Concentrations of all species are constant (not changing)
3. The system must be **closed** (no matter enters or leaves)

One thing that is NOT true: concentrations are NOT necessarily equal — reactants and products can be present in very different amounts at equilibrium.

EXAM ALERT

"Dynamic" means both reactions are still happening. This is a common exam trap — students write that equilibrium means "the reaction has stopped." It has NOT stopped. Both forward and reverse reactions continue, but at equal rates.

Le Chatelier's Principle

Le Chatelier's Principle: When a system at equilibrium is subjected to a change (a "disturbance" or "stress"), the system will respond by shifting in the direction that **opposes** the change, in order to re-establish equilibrium.

The "shift" means the rate of one direction temporarily exceeds the other until a new equilibrium is established.

Applying Le Chatelier's Principle

Consider the general equilibrium:

$$\mathrm{A(g)} + \mathrm{B(g)} \rightleftharpoons \mathrm{C(g)} + \mathrm{D(g)} \quad \Delta H < 0 \text{ (exothermic)}$$

Change 1: Concentration

Disturbance	System Response	Direction of Shift
Add more reactant $[\mathrm{A}]$ or $[\mathrm{B}]$	Equilibrium shifts to reduce added reactant; produces more C and D	Shift right (forward)
Remove reactant $[\mathrm{A}]$ or $[\mathrm{B}]$	Equilibrium shifts to replace removed reactant	Shift left (reverse)
Add more product $[\mathrm{C}]$ or $[\mathrm{D}]$	Equilibrium shifts to reduce added product; consumes C and D	Shift left (reverse)
Remove product $[\mathrm{C}]$ or $[\mathrm{D}]$	Equilibrium shifts to replace removed product; makes more product	Shift right (forward)

Change 2: Temperature

Temperature changes are unique because they actually change the value of K_c (unlike concentration and pressure, which only shift the position).

Treat heat as a "reactant" or "product":

- For an **exothermic** forward reaction: heat is a product. Increasing temperature is like adding a product → shift left (reverse).
- For an **endothermic** forward reaction: heat is a reactant. Increasing temperature is like adding a reactant → shift right (forward).

Change	Exothermic forward ($\Delta H < 0$)	Endothermic forward ($\Delta H > 0$)
Increase temperature	Shift left; K_c decreases	Shift right; K_c increases
Decrease temperature	Shift right; K_c increases	Shift left; K_c decreases

Change 3: Pressure (gases only)

Changing pressure has an effect only if there are **different numbers of moles of gas** on each side of the equation.

- Increasing pressure → system shifts toward the side with **fewer moles of gas** (to reduce pressure)
- Decreasing pressure → system shifts toward the side with **more moles of gas** (to increase pressure)
- If moles of gas are equal on both sides → pressure change has **no effect** on equilibrium position

Change 4: Catalyst

Adding a catalyst speeds up **both** the forward and reverse reactions **equally**. Therefore:

- The equilibrium position does **not** shift (no change in concentrations at equilibrium)
- K_c is **unchanged**
- The system reaches equilibrium **faster** — that is the only effect

Summary Table — Le Chatelier Disturbances

Disturbance	Direction of Shift	Effect on K_c
Add reactant	Right (forward)	No change
Remove reactant	Left (reverse)	No change
Add product	Left (reverse)	No change
Remove product	Right (forward)	No change
Increase temperature (exothermic forward)	Left (reverse)	Decreases
Increase temperature (endothermic forward)	Right (forward)	Increases
Decrease temperature (exothermic forward)	Right (forward)	Increases
Increase pressure (fewer gas moles on right)	Right (forward)	No change
Increase pressure (fewer gas moles on left)	Left (reverse)	No change
Increase pressure (equal gas moles both sides)	No shift	No change
Add catalyst	No shift	No change

⚠ **EXAM ALERT**

Pressure vs concentration: Adding an inert gas at constant volume does NOT shift the equilibrium — it does not change the partial pressures of the reactants or products. Adding an inert gas at constant pressure DOES dilute all species, effectively decreasing concentrations, but IB SL questions rarely go this far. When in doubt, state whether total pressure or partial pressure of reactants is changing.

WORKED EXAMPLE

Worked Example 5.1 — Le Chatelier's Principle

Consider the equilibrium:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \quad \Delta H = -197 \text{ kJ mol}^{-1}$$

Predict the direction of shift for each change:

(a) Temperature is increased from 500 °C to 600 °C.

(b) Pressure is doubled.

(c) Some SO_3 is removed from the equilibrium mixture.

(d) A vanadium(V) oxide (V_2O_5) catalyst is added.

Solutions:

(a) The forward reaction is exothermic ($\Delta H = -197 \text{ kJ mol}^{-1}$). Increasing temperature is equivalent to adding heat (a product in the exothermic direction). The system opposes this by shifting **left** (reverse), decreasing $[SO_3]$ and increasing $[SO_2]$ and $[O_2]$. K_c decreases.

(b) Left side: 2 + 1 = **3 moles of gas**. Right side: **2 moles of gas**. Increasing pressure favours the side with fewer moles of gas. Shift **right** (forward), producing more SO_3. K_c unchanged.

(c) Removing product SO_3 reduces $[SO_3]$. System shifts **right** (forward) to replace the removed product, consuming SO_2 and O_2. K_c unchanged.

(d) The catalyst (V_2O_5) speeds up both forward and reverse reactions equally. **No shift.** Equilibrium is reached faster, but the equilibrium position and K_c are unchanged.

6. Equilibrium Constant K_c

Writing the K_c Expression

For any reversible reaction at equilibrium:

$$a\text{A} + b\text{B} \rightleftharpoons c\text{C} + d\text{D}$$

The equilibrium constant K_c is defined as:

$$K_c = \frac{[\text{C}]^c[\text{D}]^d}{[\text{A}]^a[\text{B}]^b}$$

Where:

- Square brackets $[\,]$ denote equilibrium concentration in mol dm^{-3}

- The **products** go in the numerator (top)
- The **reactants** go in the denominator (bottom)
- Each concentration is raised to the power of its **stoichiometric coefficient** from the balanced equation

MEMORISE THIS

The K_c expression rule:

$$K_c = \frac{\text{products (concentration)}^{\text{coefficient}}}{\text{reactants (concentration)}^{\text{coefficient}}}$$

Pure solids and pure liquids are NOT included in the K_c expression — their concentrations are constant and are incorporated into the value of K_c itself.

Examples:

- $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$: $K_c = [CO_2]$ (solids omitted)
- $CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$: water (pure liquid) omitted

Magnitude of K_c

Value of K_c	Meaning
$K_c \gg 1$ (e.g. $K_c = 10^{10}$)	Products strongly favoured; reaction goes nearly to completion
$K_c \approx 1$	Roughly equal amounts of reactants and products at equilibrium
$K_c \ll 1$ (e.g. $K_c = 10^{-10}$)	Reactants strongly favoured; very little product formed

What Changes K_c?

K_c depends **only on temperature**. Changing concentration, pressure, or adding a catalyst does NOT change K_c — only the position of equilibrium shifts.

EXAM ALERT

This is one of the most commonly tested IB points: “Which of the following changes will alter the value of K_c?” The answer is always **temperature only**. Concentration, pressure, and catalysts change the rate or shift the position of equilibrium, but K_c stays the same.

Worked Example — Calculating K_c

WORKED EXAMPLE

Worked Example 6.1 — K_c from Equilibrium Concentrations

At 500 K, the following equilibrium is established:

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

At equilibrium: $[H_2] = 0.10$ mol dm^{-3}, $[I_2] = 0.10$ mol dm^{-3}, $[HI] = 0.74$ mol dm $^{-3}$

Calculate K_c.

Solution:

Write the K_c expression:

$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

Substitute equilibrium concentrations:

$$K_c = \frac{(0.74)^2}{(0.10)(0.10)} = \frac{0.5476}{0.0100} = 54.8$$

$K_c = 54.8$ (no units for this equilibrium since the powers cancel: $\frac{\text{mol}^2\ \text{dm}^{-6}}{\text{mol}^2\ \text{dm}^{-6}}$)

Interpretation: $K_c = 54.8 \gg 1$, so products (HI) are favoured at this temperature. Most of the hydrogen and iodine has been converted to hydrogen iodide at equilibrium.

Worked Example — ICE Table

WORKED EXAMPLE

Worked Example 6.2 — ICE Table to Find Equilibrium Concentration

At a certain temperature, $K_c = 4.0$ for the reaction:

$$\mathrm{A(g)} + \mathrm{B(g)} \rightleftharpoons \mathrm{C(g)} + \mathrm{D(g)}$$

0.80 mol of A and 0.80 mol of B are placed in a 1.00 dm^3 flask. Calculate the equilibrium concentrations of all species.

Solution — ICE Table:

The ICE table tracks Initial concentrations, Change, and Equilibrium concentrations:

Species	[A]	[B]	[C]	[D]
Initial / mol dm^{-3}	0.80	0.80	0	0
Change / mol dm^{-3}	$-x$	$-x$	$+x$	$+x$
Equilibrium / mol dm^{-3}	$0.80 - x$	$0.80 - x$	x	x

Write the K_c expression and substitute:

$$K_c = \frac{[\mathrm{C}][\mathrm{D}]}{[\mathrm{A}][\mathrm{B}]} = \frac{(x)(x)}{(0.80-x)(0.80-x)} = \frac{x^2}{(0.80-x)^2} = 4.0$$

Take the square root of both sides:

$$\frac{x}{0.80-x} = \sqrt{4.0} = 2.0$$

Solve for x:

$$x = 2.0(0.80 - x) = 1.60 - 2.0x$$

$$3.0x = 1.60$$

$$x = 0.533 \text{ mol dm}^{-3}$$

Equilibrium concentrations:

- $[\mathrm{A}] = [\mathrm{B}] = 0.80 - 0.533 = \mathbf{0.27}$ mol dm^{-3}
- $[\mathrm{C}] = [\mathrm{D}] = x = \mathbf{0.53}$ mol dm^{-3}

Check: $K_c = \frac{(0.533)^2}{(0.267)^2} = \frac{0.284}{0.0713} \approx 4.0$ ✓

7. Industrial Application — The Haber Process

The Reaction

The Haber Process is the industrial synthesis of ammonia from nitrogen and hydrogen:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92 \text{ kJ mol}^{-1}$

Ammonia is one of the most important industrial chemicals — it is the starting material for fertilisers, explosives, cleaning products, and many other chemicals. Understanding the Haber Process means applying both Le Chatelier's Principle and collision theory to choose the best industrial conditions.

Applying Le Chatelier's Principle

Pressure

- **Left side:** $1 + 3 =$ **4** moles of gas
- **Right side: 2** moles of gas
- **High pressure** favours the right side (fewer moles of gas) → more NH_3 produced
- **Industrial pressure used: ~200 atm**

However, very high pressures are extremely expensive to maintain (requires thick-walled vessels, powerful compressors, high energy costs) and pose safety risks. ~200 atm is a compromise between yield and cost.

Temperature

- The forward reaction is **exothermic** ($\Delta H = -92 \text{ kJ mol}^{-1}$)
- **Low temperature** shifts equilibrium right → higher yield of NH_3
- But low temperature means a **very slow rate** of reaction (molecules have less energy; fewer exceed E_a)
- **High temperature** means faster rate but lower equilibrium yield
- **Industrial temperature used: 400–500 °C**

This is a critical compromise: at 400–500 °C, the rate is fast enough to be economically viable, even though the equilibrium yield of NH_3 is only ~15%. The ammonia is continuously removed (liquefied and extracted), which shifts equilibrium right (Le Chatelier) and allows the unreacted N_2 and H_2 to be recycled.

Catalyst

- An **iron catalyst** (with potassium oxide and alumina as promoters) is used
- The catalyst increases the rate of reaction without affecting K_c or the equilibrium position
- This allows the reaction to proceed at an acceptable rate at the lower 400–500 °C temperature

Summary of Haber Process Conditions

Condition	Value	Reason
Pressure	~200 atm	High pressure favours products (fewer gas moles); limited by cost/safety
Temperature	400–500 °C	Compromise: fast enough rate; acceptable (but not maximum) yield
Catalyst	Iron (+ promoters)	Increases rate; does not change equilibrium position or K_c
NH_3 **removal**	Continuous liquefaction	Removes product; shifts equilibrium right by Le Chatelier; unreacted gases recycled

EXAM ALERT

The Haber Process is one of the most frequently examined topics in IB Chemistry SL. A typical 4–6 mark question asks you to explain the choice of conditions. You must:

1. State the condition used (e.g. “400–500 °C”)
2. Explain the equilibrium reasoning (e.g. “lower temperature would give higher yield because the forward reaction is exothermic, but…”)
3. Explain the rate reasoning (e.g. “too low a temperature would make the rate impractically slow”)
4. Use the word **compromise** — the IB mark scheme almost always awards a mark for this

For pressure, mention both the yield benefit (fewer moles of gas on right) AND the practical cost/safety limitation.

IB TIP

The K_c expression for the Haber Process is:

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

Notice $[H_2]$ is cubed because the stoichiometric coefficient of H_2 is 3. This expression comes up frequently in data-response questions.

WORKED EXAMPLE

Worked Example 7.1 — Haber Process K_c Calculation

At 500 °C, the equilibrium concentrations in a Haber Process reactor are:

$[N_2] = 3.0 \text{ mol dm}^{-3}$, $[H_2] = 9.0 \text{ mol dm}^{-3}$, $[NH_3] = 2.4 \text{ mol dm}^{-3}$

Calculate K_c at this temperature.

Solution:

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

$$K_c = \frac{(2.4)^2}{(3.0)(9.0)^3} = \frac{5.76}{(3.0)(729)} = \frac{5.76}{2187} = 2.63 \times 10^{-3}$$

$K_c = 2.63 \times 10^{-3}$ at 500 °C

Interpretation: $K_c \ll 1$, confirming that at 500 °C, the equilibrium lies to the left — at this temperature, only a small fraction of N_2 and H_2 is converted to NH_3 at equilibrium (~15% yield). This is why the Haber Process recycles unreacted gases.

8. Practice MCQs

The following 10 questions are written in IB exam style. Try each question before revealing the answer.

▶ **MCQ 1.** Which of the following changes will increase the rate of a reaction between zinc and dilute hydrochloric acid? I. Using more concentrated acid II. Increasing the surface area of the zinc III. Adding a catalyst A. I only B. I and II only C. II and III only D. I, II, and III

▶ **MCQ 2.** A reaction is carried out at two temperatures: 30 °C and 50 °C. Using the Maxwell-Boltzmann distribution, which statement best explains the increased rate at 50 °C? A. The activation energy decreases at higher temperature B. More molecules have energy greater than or equal to the activation energy C. The molecules collide more frequently only D. The activation energy is exceeded by all molecules at 50 °C

▶ **MCQ 3.** For the equilibrium: $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$ What is the effect of increasing pressure on the equilibrium position? A. Shifts right; more PCl_3 and Cl_2 are produced B. Shifts left; more PCl_5 is produced C. No shift; pressure does not affect gas-phase equilibria D. Shifts right; K_c increases

▶ **MCQ 4.** Which of the following will change the value of K_c for a reaction at equilibrium? A. Adding more reactant B. Increasing the pressure C. Adding a catalyst D. Raising the temperature

▶ **MCQ 5.** Consider the reaction: $2NO_2(g) \rightleftharpoons N_2O_4(g) \quad \Delta H = -57 \text{ kJ mol}^{-1}$ What are the effects of simultaneously increasing temperature and decreasing pressure? A. Both changes shift the equilibrium to the right B. Temperature shifts left; pressure shifts right — net effect depends on magnitudes C. Temperature shifts right; pressure shifts left D. Temperature shifts left; pressure shifts left

▶ **MCQ 6.** At equilibrium, $K_c = 1.2 \times 10^{-5}$ for a reaction. Which statement is correct? A. The forward reaction is faster than the reverse reaction B. The reaction has not yet reached equilibrium C. Reactants are strongly favoured at equilibrium D. Products are strongly favoured at equilibrium

▶ **MCQ 7.** Which expression correctly gives K_c for: $\mathrm{N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)}$ A. $K_c = \frac{[\mathrm{NH_3}]}{[\mathrm{N_2}][\mathrm{H_2}]}$ B. $K_c = \frac{[\mathrm{NH_3}]^2}{[\mathrm{N_2}][\mathrm{H_2}]^3}$ C. $K_c = \frac{[\mathrm{N_2}][\mathrm{H_2}]^3}{[\mathrm{NH_3}]^2}$ D. $K_c = \frac{2[\mathrm{NH_3}]}{[\mathrm{N_2}] + 3[\mathrm{H_2}]}$

▶ **MCQ 8.** The Haber Process uses a temperature of 400–500 °C rather than a lower temperature such as 100 °C. Why? A. A higher temperature increases K_c, favouring product formation B. The forward reaction is endothermic, so higher temperature increases yield C. A lower temperature would give a higher yield but an unacceptably slow rate D. The catalyst only works above 400 °C

▶ **MCQ 9.** Which of the following correctly describes a heterogeneous catalyst? A. The catalyst and reactants are in the same physical state B. The catalyst is consumed during the reaction and must be regenerated C. The catalyst changes the enthalpy change (ΔH) of the reaction D. The catalyst is in a different physical state from the reactants

▶ **MCQ 10.** For the equilibrium: $\mathrm{Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons [FeSCN]^{2+}(aq)}$ The product $\mathrm{[FeSCN]^{2+}}$ is deep red. A solution at equilibrium is pale orange. What would be observed if excess $\mathrm{Fe^{3+}}$ ions are added? A. The solution becomes paler B. The solution becomes deeper red C. No change — the system is already at equilibrium D. The solution turns colourless

Summary — Key Formulas and Definitions

Concept	Formula / Key Statement
Rate of reaction	$\text{Rate} = \frac{\Delta[\text{concentration}]}{\Delta t}$ in mol dm^{-3} s^{-1}
Effective collision	Requires sufficient energy (≥ E_a) AND correct orientation
Activation energy (E_a**)**	Minimum energy for reaction to occur
Catalyst effect	Provides lower-E_a alternative pathway; does not change ΔH or K_c
Dynamic equilibrium	Forward rate = reverse rate; concentrations constant; system is closed
Le Chatelier's Principle	System opposes any imposed change; shifts to re-establish equilibrium
K_c **expression**	$K_c = \frac{[\text{products}]^{\text{coeff}}}{[\text{reactants}]^{\text{coeff}}}$ (exclude pure solids/liquids)
K_c **and temperature**	Only temperature changes K_c
Haber Process	$\mathrm{N_2 + 3H_2 \rightleftharpoons 2NH_3}$, $\Delta H = -92$ kJ mol^{-1}; 400–500 °C, ~200 atm, Fe catalyst

IB TIP

In the IB exam, the most marks are lost on:

1. Not explaining collision theory fully (saying “faster” without saying “more effective collisions” or linking to E_a)
2. Confusing “equilibrium position shifts” with ”K_c changes” — only temperature changes K_c
3. Forgetting to count gas moles correctly before applying pressure Le Chatelier reasoning
4. Writing K_c expressions with addition instead of multiplication, or with incorrect coefficients

May 2026 Prediction Questions

⚠ EXAM ALERT

These are NOT official IB questions. These are trend-based practice questions written to reflect the topic areas and question styles most likely to appear on the May 2026 IB Chemistry SL Paper 2. Based on recent exam patterns (2022-2025), expect heavy weighting on: collision theory and factors affecting rate, K_c expressions and calculations, and Le Chatelier’s principle applied to industrial processes.

WORKED EXAMPLE

Question 1 [Collision Theory and Rate Factors] [~6 marks]

The reaction between hydrochloric acid and magnesium ribbon is studied:

$\mathrm{Mg}(s) + 2\mathrm{HCl}(aq) \rightarrow \mathrm{MgCl_2}(aq) + \mathrm{H_2}(g)$

(a) Using collision theory, explain why increasing the concentration of HCl increases the rate of this reaction.

(b) Explain why increasing the temperature increases the rate of reaction. Your answer should refer to the Maxwell-Boltzmann distribution.

(c) A catalyst is added. Explain how a catalyst increases the rate without being consumed.

▶ Show Solution

WORKED EXAMPLE

Question 2 [K_c Calculation] [~7 marks]

The equilibrium reaction below was studied at 450 $^\circ\mathrm{C}$:

$\mathrm{H_2}(g) + \mathrm{I_2}(g) \rightleftharpoons 2\mathrm{HI}(g)$

At equilibrium, the concentrations in a $1.0\ \mathrm{dm^3}$ container are: $[\mathrm{H_2}] = 0.11\ \mathrm{mol\ dm^{-3}}$, $[\mathrm{I_2}] = 0.11\ \mathrm{mol\ dm^{-3}}$, $[\mathrm{HI}] = 0.78\ \mathrm{mol\ dm^{-3}}$.

(a) Write the expression for K_c for this reaction.

(b) Calculate the value of K_c at 450 $^\circ\mathrm{C}$.

(c) If the temperature is increased and K_c decreases, deduce whether the forward reaction is exothermic or endothermic. Explain your reasoning.

▶ Show Solution

WORKED EXAMPLE

Question 3 [Le Chatelier's Principle] [~7 marks]

The Haber process for the manufacture of ammonia is represented by:

$\mathrm{N_2}(g) + 3\mathrm{H_2}(g) \rightleftharpoons 2\mathrm{NH_3}(g) \quad \Delta H = -92\ \mathrm{kJ\ mol^{-1}}$

Industrial conditions: 400-500 $^\circ\mathrm{C}$, ~200 atm, iron catalyst.

(a) Use Le Chatelier's principle to predict and explain the effect of increasing pressure on the yield of ammonia.

(b) The forward reaction is exothermic. According to Le Chatelier's principle, a low temperature would favour ammonia production. Explain why the industrial process uses 400-500 $^\circ\mathrm{C}$ despite this.

(c) Explain why the iron catalyst does not affect the position of equilibrium or the value of K_c.

▶ Show Solution