MEGA LECTURE

IB DP CHEMISTRY

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

Reactivity 2.3 How Far? The Extent of Chemical Change

Revision Notes · Standard and Higher Level

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Original notes prepared for the IB Diploma Programme Chemistry course (2025 syllabus)

What the syllabus requires

Reactivity 2.3 answers the question *how far* a reaction proceeds before it appears to stop. Many reactions do not go to completion; instead they settle at a point of balance called equilibrium. Use the list below as a final checklist.

Understanding	You should be able to...
Dynamic equilibrium	Describe the features of a reversible reaction that has reached equilibrium in a closed system.
Equilibrium constant K_c	Deduce the K_c expression from a balanced equation and interpret the size of K_c.
Reaction quotient Q	Compare Q with K_c to predict the direction in which a system will shift.
Le Châtelier's principle	Predict and explain the effect of concentration, pressure, temperature and catalysts on the position of equilibrium.
What changes K	Recognise that only a change of temperature alters the value of K_c.
Industrial processes	Explain the compromise conditions used in the Haber and Contact processes.
(HL) K_c calculations	Use ICE tables to calculate K_c or equilibrium concentrations, including the small-K approximation.
(HL) ΔG° and K	Relate the standard Gibbs energy change to K through $\Delta G^\circ = -RT \ln K$.

> **Exam note:** Sections 1-6 are common to SL and HL. Sections marked **(HL)** — the K_c calculations and the link to Gibbs energy — are assessed at Higher Level only, but every student benefits from reading them.

1. Dynamic equilibrium

A **reversible reaction** is one that can proceed in both the forward and the reverse direction under the same conditions. We write it with a double half-arrow, for example the formation of hydrogen iodide:

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

When such a reaction is carried out in a **closed system** (no matter enters or leaves), it eventually reaches a state of **dynamic equilibrium**. The word *dynamic* is essential: the reaction has not stopped. Both the forward and reverse reactions continue, but they now occur at exactly the same rate, so there is no net change.

Features of a system at equilibrium

- The system is **closed** — no reactants or products escape.
- The **rate of the forward reaction equals the rate of the reverse reaction**.
- **Macroscopic properties are constant** — concentration, colour, pressure and density no longer change with time.
- Both reactions **continue at the molecular level**; equilibrium is dynamic, not static.
- Equilibrium can be approached from **either direction** and gives the same final mixture for the same conditions.

Physical and chemical equilibria

Equilibrium is not confined to chemical reactions. A **physical equilibrium** involves a change of state or dissolving, with no new substance formed. In a sealed bottle of water, for example, molecules evaporate and condense at equal rates:

$$H_2O(l) \rightleftharpoons H_2O(g)$$

The dissolving of a sparingly soluble salt, or of carbon dioxide in a fizzy drink, is also a physical equilibrium. A **chemical equilibrium** involves the making and breaking of bonds, as in the Haber process or the dissociation of a weak acid. In both cases the same underlying idea applies: two opposing processes running at equal rates.

How the system reaches equilibrium

Picture starting with only reactants. At the very beginning the forward rate is at its maximum (reactant concentration is highest) while the reverse rate is zero (no product yet). As the reaction proceeds, reactants are used up so the forward rate **falls**, and product builds up so the reverse rate **rises**. Equilibrium is reached at the instant the two rates become equal — from then on the concentrations stay constant.

Stage	Forward rate	Reverse rate	Concentrations
Start (reactants only)	Maximum	Zero	Reactants high, products zero
During approach	Falling	Rising	Reactants → down, products → up
At equilibrium	Constant	Equal to forward	All constant (not necessarily equal)

On a graph of rate against time the two curves meet and then run together horizontally; on a graph of concentration against time the curves level off to constant plateau values. The plateau does not mean the reaction has stopped — it means the opposing rates have become equal.

> **Key phrase:** If asked to define equilibrium, always include three ideas — closed system, equal forward and reverse rates, and constant macroscopic properties. Dropping the word “dynamic” is a common way to lose a mark.

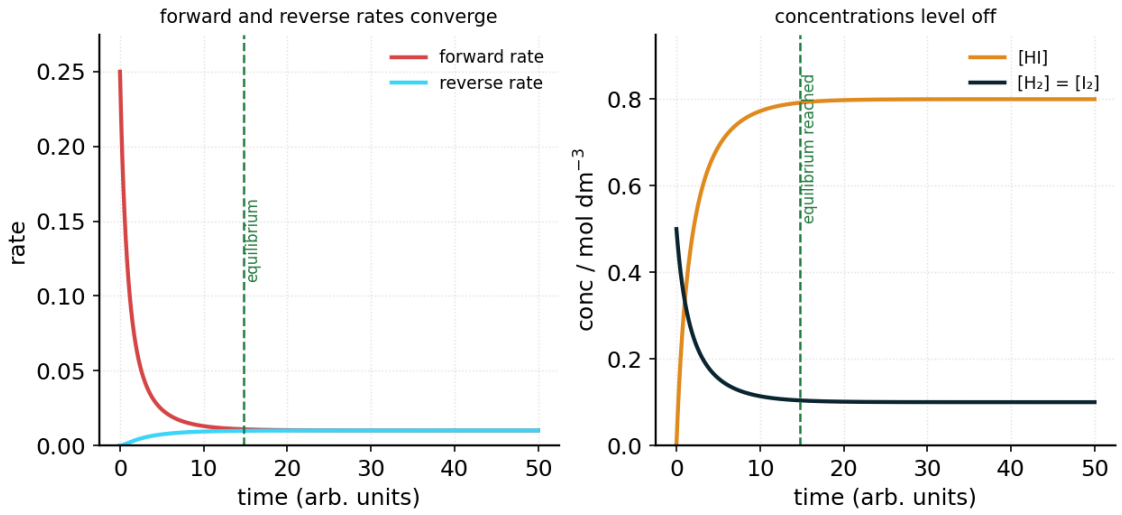

Figure 1. Approach to equilibrium for $H_2 + I_2 \rightleftharpoons 2HI$. Left: the forward rate falls and the reverse rate rises until they are equal (dashed line). Right: the concentrations level off at constant, non−zero values, [HI] = 0.80 and $[H_2] = [I_2]$ = 0.10 mol dm^{-3} (K_c = 64).

2. The equilibrium constant, K_c

Although the concentrations of reactants and products are constant at equilibrium, they are not usually equal to one another. For a given temperature, however, a fixed mathematical relationship always holds between them. This is captured by the **equilibrium constant**, K_c.

Writing the expression

For the general reaction

$$\mathbf{aA + bB \rightleftharpoons cC + dD}$$

the equilibrium constant in terms of concentration is

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

The rules are simple and worth memorising:

- **Products** (right-hand side) go on **top**; **reactants** go on the **bottom**.
- Each concentration is raised to a **power equal to its coefficient** in the balanced equation.
- Square brackets [ ] mean "equilibrium concentration in mol dm^{-3}".
- **Pure solids and pure liquids are omitted** — their concentration (density) is effectively constant and is absorbed into K_c.

So for the decomposition of calcium carbonate, only the gas appears:

$$\mathbf{CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)\ K_c = [CO_2]}$$

For gaseous equilibria we may instead use partial pressures, giving $\mathbf{K_p}$, written the same way but with the equilibrium partial pressure of each gas in place of its concentration. Pure solids and liquids are again left out.

The meaning of the magnitude of K

The size of K is a direct measure of *how far* a reaction proceeds — of the extent of reaction — at a given temperature.

Value of K	Interpretation	Position of equilibrium
K > 1 (large)	Numerator dominates: products are favoured. A very large K (e.g. 10^{10}) means the reaction is essentially complete.	Lies to the right (products)
K ≈ 1	Comparable amounts of reactants and products present at equilibrium.	Roughly central
K < 1 (small)	Denominator dominates: reactants are favoured. A very small K (e.g. 10^{-10}) means little reaction occurs.	Lies to the left (reactants)

> **Important:** K tells you **how far** a reaction goes, not **how fast**. A reaction can have a huge K yet be immeasurably slow (rate is Reactivity 2.2). Extent and rate are separate ideas.

Worked example 1 — writing K_c expressions

Write the K_c expression for each equilibrium.

(a) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

$K_c = [NH_3]^2 / ([N_2][H_2]^3)$ — note the power 3 on H_2 from its coefficient.

(b) $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$

$K_c = [CH_3COO^-][H^+] / [CH_3COOH]$. (This particular K_c is the acid dissociation constant, K_a.)

(c) $C(s) + CO_2(g) \rightleftharpoons 2CO(g)$

$K_c = [CO]^2 / [CO_2]$. The pure solid carbon is omitted.

3. The reaction quotient, Q

The **reaction quotient**, Q, has exactly the same algebraic form as K_c, but it is calculated using the concentrations present at *any* moment, not necessarily at equilibrium. Comparing Q with K tells us which way a reaction must shift to reach equilibrium.

Comparison	Meaning	Which way does the reaction go?
Q < K	Too few products relative to equilibrium	Net forward reaction (→ products) until Q rises to K
Q = K	System is already at equilibrium	No net change
Q > K	Too many products relative to equilibrium	Net reverse reaction (→ reactants) until Q falls to K

A helpful mental picture: the reaction always shifts in the direction that brings Q closer to K. If you disturb an equilibrium (Section 4), the disturbance changes Q away from K, and the system responds by moving to restore Q = K.

Worked example 2 — predicting direction with Q

For $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$, $K_c = 64$. A mixture is prepared with $[H_2] = 0.30$, $[I_2] = 0.30$ and $[HI] = 0.60$ mol dm^{-3}. In which direction does the reaction proceed?

$Q = [HI]^2 / ([H_2][I_2]) = (0.60)^2 / (0.30 \times 0.30) = 0.36 / 0.09 = 4.0$.

Q (4.0) < K (64), so there are too few products. The reaction proceeds **forward** (→ more HI) until Q rises to 64.

4. Le Châtelier's principle

Le Châtelier's principle states that if a system at equilibrium is subjected to a change, the position of equilibrium shifts in the direction that **opposes** (partially counteracts) that change. It lets us predict the direction of a shift without any calculation. We illustrate each factor with the exothermic gas-phase equilibrium

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

Changing a concentration

Increasing the concentration of a reactant makes Q < K, so the system shifts **forward** to consume the added reactant and make more product. Removing a product has the same effect. Conversely, adding product or removing reactant shifts the equilibrium **backward**. Continuously removing a product is a common industrial trick to keep pulling a reaction forward.

Changing pressure or volume (gases)

Pressure matters only when the number of moles of **gas** differs between the two sides. Increasing the pressure (by decreasing the volume) shifts the equilibrium towards the side with **fewer moles of gas**, which reduces the pressure again. In the ammonia equilibrium there are 4 mol of gas on the left and 2 mol on the right, so higher pressure favours the **forward** reaction and more ammonia. If the moles of gas are equal on both sides, pressure has no effect on the position.

Changing temperature

Temperature is the one factor that shifts equilibrium **and** changes K. Treat heat as a "reactant" or "product" using the sign of ΔH:

- For an **exothermic** forward reaction (ΔH negative), heat is released. **Raising** the temperature shifts equilibrium **backward** (endothermic direction), lowering the yield and decreasing K.
- **Lowering** the temperature of an exothermic reaction shifts it **forward**, raising the yield and increasing K.
- For an **endothermic** forward reaction (ΔH positive), the effects are reversed: heating increases yield and K.

Adding a catalyst

A catalyst speeds up the forward and reverse reactions **equally**. It therefore **has no effect on the position of equilibrium or on the value of K**; it only allows equilibrium to be **reached more quickly**. This is why catalysts are vital industrially — they save time and energy without changing the maximum possible yield.

Change imposed	Direction of shift ($N_2 + 3H_2 \rightleftharpoons 2NH_3$, exothermic)	Does K change?
Add N_2 or H_2	Forward (makes more NH_3)	No
Remove NH_3	Forward	No
Increase pressure	Forward (towards 2 mol side)	No
Increase temperature	Backward (endothermic direction)	Yes — K decreases
Add catalyst	No shift (equilibrium reached faster)	No

> **Exam technique:** Always give a **reason** with the direction. For pressure, state the moles of gas on each side; for temperature, quote the sign of ΔH and identify the endothermic direction.

A visible equilibrium: the cobalt complex

Colour changes make equilibrium shifts easy to observe. A classic example is the equilibrium between the pink hydrated cobalt(II) ion and the blue chloro-complex:

$$[Co(H_2O)_6]^{2+}(aq) + 4Cl^-(aq) \rightleftharpoons [CoCl_4]^{2-}(aq) + 6H_2O(l) \quad \Delta H \text{ positive}$$

The pink form is on the left, the blue form on the right. Le Châtelier predicts the colour change for each disturbance:

Change	Shift	Observation
Add concentrated HCl (raises $[Cl^-]$)	Forward	Solution turns blue
Add water (dilutes Cl^-)	Backward	Solution turns pink
Warm the mixture	Forward (endothermic direction)	Turns blue; K increases with T
Cool the mixture	Backward	Turns pink; K decreases

The temperature results confirm Section 5: because this reaction is endothermic, heating shifts it towards products and increases K, while cooling does the reverse. Adding or diluting chloride ions shifts the position but leaves K unchanged.

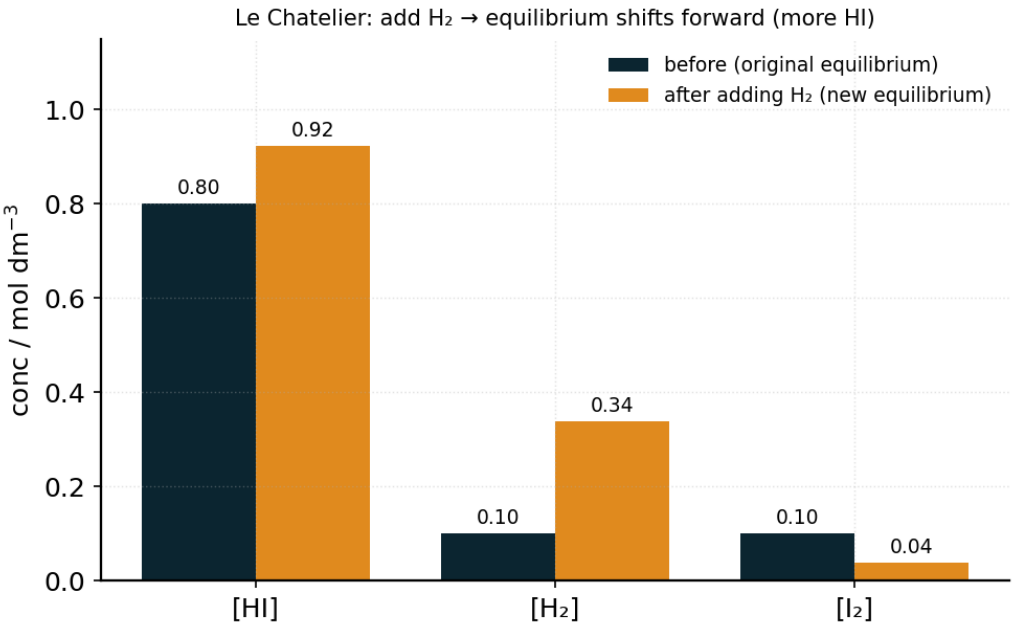

Figure 3. Le Chatelier's principle (concentration change). Adding H_2 to the equilibrium $H_2 + I_2 \rightleftharpoons 2HI$ ($K_c = 64$) makes $Q < K$, so the position shifts forward: [HI] rises from 0.80 to 0.92 mol dm^{-3} as I_2 is consumed, yet K_c itself is unchanged.

5. What actually changes the value of K

A frequent misunderstanding is that anything which shifts the equilibrium must also change K. It does not. K depends on **temperature alone**.

- Changing **concentration** shifts the position but the system returns to the **same K** — concentrations rearrange so that the ratio is restored.
- Changing **pressure or volume** shifts the position but leaves K unchanged.
- Adding a **catalyst** changes neither the position nor K.
- Only a change of **temperature** changes K.

The reason follows from Section 4. When you raise the temperature of an exothermic reaction, the equilibrium moves back towards reactants; the new equilibrium mixture has proportionally more reactants and fewer products, so the ratio [products]/[reactants] — that is, K — is genuinely smaller. For an

endothermic reaction the reverse is true and K grows with temperature. Because concentration and pressure changes only rearrange the existing species, the ratio always returns to the same value once equilibrium is re-established.

6. Industrial applications

Industrial chemists must satisfy three competing demands: a high **yield** (favoured by equilibrium position), a fast **rate** (favoured by kinetics) and a low **cost**. The chosen conditions are always a **compromise**, and examiners love to ask you to justify them.

The Haber process

Ammonia is made for fertilisers by combining nitrogen and hydrogen:

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

Condition	What equilibrium/rate wants	Compromise chosen	Reason
Pressure	High pressure raises yield (4 mol → 2 mol) and rate	≈ 200 atm	Very high pressure is costly and dangerous (thick pipes, safety)
Temperature	Low temperature raises yield (exothermic) but is slow	≈ 450 °C	A moderate temperature gives an acceptable rate at a workable yield
Catalyst	Speeds up equilibrium without changing yield	Finely divided iron	Cheap catalyst gives an economic rate; no effect on position

Unreacted N_2 and H_2 are recycled, and ammonia is removed by cooling and liquefying it. Removing the product continually shifts the equilibrium forward, improving the overall conversion despite the modest single-pass yield.

Worked example 3 — reasoning about Haber conditions

Explain, in terms of equilibrium and rate, why the Haber process does **not** use a very low temperature even though this would maximise the yield of ammonia.

The forward reaction is exothermic ($\Delta H = -92\ kJ\ mol^{-1}$). By Le Châtelier, lowering the temperature shifts the equilibrium forward and increases both the yield and the value of K — good for extent.

However, at a low temperature the **rate** is far too slow, because few molecules have enough energy to react. Ammonia would be produced too slowly to be economic.

A moderate temperature of about 450 °C is therefore chosen as a **compromise**: it sacrifices some equilibrium yield in return for a commercially acceptable rate. An iron catalyst and continual removal of NH_3 further improve the overall output.

The Contact process

Sulfuric acid manufacture depends on the reversible oxidation of sulfur dioxide, the key equilibrium step:

$$\mathbf{2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)\ \Delta H = -196\ kJ\ mol^{-1}}$$

Condition	What equilibrium/rate wants	Compromise chosen	Reason
Pressure	High pressure raises yield (3 mol → 2 mol)	≈ 1-2 atm (near atmospheric)	Yield is already ~98%, so costly high pressure is not worth it
Temperature	Low temperature raises yield (exothermic) but is slow	≈ 450 °C	Balances a high yield against an acceptable rate
Catalyst	Speeds up equilibrium; no effect on yield	Vanadium(V) oxide, V_2O_5	Provides an economic rate at moderate temperature

> **Compromise summary:** In both processes a **moderate temperature** sacrifices some equilibrium yield to gain a workable rate, a **catalyst** speeds things up without cost to yield, and **pressure** is set as high as economics safely allow. Learn to state the trade-off, not just the numbers.

7. (HL) Calculations with K_c

At Higher Level you must calculate K_c from equilibrium data, and find equilibrium concentrations from a known K_c. The organising tool is the **ICE table** — **I**nitial, **C**hange, **E**quilibrium.

Setting up an ICE table

- **I**: list the initial concentration (or moles) of every species.
- **C**: express the change using a single unknown *x*, scaled by the stoichiometric coefficients (reactants −, products +).
- **E**: add the initial and change rows to get the equilibrium concentrations, then substitute into the K_c expression.

Units of K_c

The units follow from the expression. Each concentration carries mol dm^{-3}, so the overall unit is (mol dm^{-3})$^{\Delta n}$ where Δn = (total power on top) − (total power on bottom). When the totals of the powers on top and bottom are equal, K_c is **dimensionless** (no units). For $H_2 + I_2 \rightleftharpoons 2HI$, top power 2 and bottom power 2 cancel, so K_c has no units.

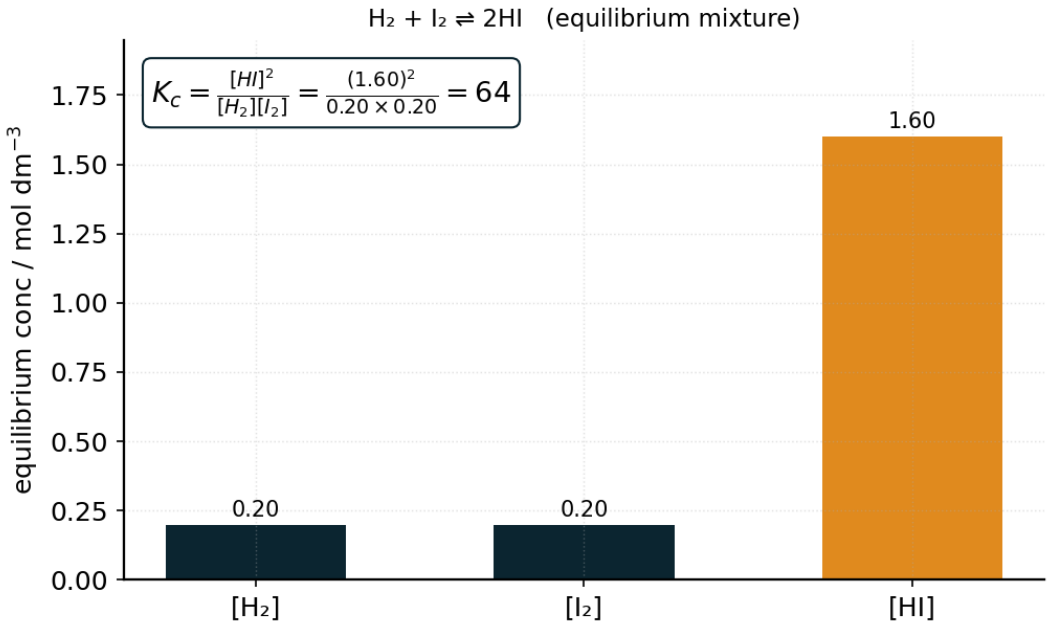

Figure 2. Worked equilibrium constant. For the equilibrium mixture $[H_2] = [I_2] = 0.20$ and $[HI] = 1.60$ mol dm^{-3}, $K_c = [HI]^2 / ([H_2][I_2]) = 64$ (dimensionless, since the powers balance).

Worked example 4 (HL) — K_c from equilibrium concentrations

In the equilibrium $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$, a vessel is found to contain $[H_2] = 0.20$, $[I_2] = 0.20$ and $[HI] = 1.60$ mol dm^{-3} at equilibrium. Calculate K_c.

$K_c = [HI]^2 / ([H_2][I_2]) = (1.60)^2 / (0.20 \times 0.20)$

$= 2.56 / 0.040 =$ **64** (no units, since the powers balance).

K is large, so at this temperature the equilibrium lies well to the right — HI is favoured.

Worked example 5 (HL) — building an ICE table

0.100 mol of H_2 and 0.100 mol of I_2 are placed in a 1.00 dm^3 flask. At equilibrium 0.160 mol of HI has formed. Find K_c.

Let the change in H_2 be $-x$. Since 0.160 mol HI forms and the ratio HI : H_2 is 2 : 1, $x = 0.080$.

ICE (mol dm^{-3}):

	H_2	I_2	2HI
Initial	0.100	0.100	0
Change	−0.080	−0.080	+0.160
Equilibrium	0.020	0.020	0.160

$K_c = (0.160)^2 / (0.020 \times 0.020) = 0.0256 / 0.00040 =$ **64**.

Worked example 6 (HL) — equilibrium concentration from K_c

For $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$, $K_c = 64$ at the working temperature. If 0.500 mol of H_2 and 0.500 mol of I_2 are mixed in a 1.00 dm^3 flask, find the equilibrium [HI].

Let x = amount of H_2 reacted. Equilibrium: $[H_2] = [I_2] = 0.500 - x$, $[HI] = 2x$.

$K_c = (2x)^2 / (0.500 - x)^2 = 64$.

Take the square root of both sides: $2x / (0.500 - x) = 8$.

$2x = 8(0.500 - x) = 4.00 - 8x$, so $10x = 4.00$ and $x = 0.400$.

$[HI] = 2x =$ **0.800 mol dm^{-3}**; and $[H_2] = [I_2] = 0.100$ mol dm^{-3}.

Worked example 7 (HL) — the small-K approximation

For the dissociation $A(g) \rightleftharpoons B(g) + C(g)$, $K_c = 1.0 \times 10^{-4}$ mol dm^{-3}. 0.10 mol of A is placed in a 1.0 dm^3 flask. Estimate [B] at equilibrium.

ICE: $[A] = 0.10 - x$, $[B] = [C] = x$. So $K_c = x^2 / (0.10 - x)$.

Because K is very small, very little A reacts, so x is much less than 0.10 and $0.10 - x \approx 0.10$.

Then $x^2 \approx K_c \times 0.10 = 1.0 \times 10^{-5}$, so $x =$ **3.2×10^{-3} mol dm^{-3}**.

Check: x is about 3% of 0.10, comfortably below the usual 5% limit, so the approximation is valid.

Approximation rule: You may write (initial $- x$) $\approx$ initial only when K is small enough that x is less than about 5% of the initial value. If it exceeds 5%, solve the full quadratic instead.

8. (HL) Gibbs energy and the equilibrium constant

Reactivity 1.4 introduced the standard Gibbs energy change, ΔG°, as the test of spontaneity. There is a direct quantitative bridge between ΔG° and the equilibrium constant K:

$$\Delta G^\circ = -RT \ln K$$

where $R = 8.31$ J K^{-1} mol^{-1} and T is in kelvin. Rearranged, $K = \exp(-\Delta G^\circ / RT)$. This single equation ties together spontaneity and extent.

Sign of ΔG°	Value of ln K	Value of K	Meaning
$\Delta G^\circ < 0$ (negative)	positive	$K > 1$	Products favoured; reaction spontaneous as written
$\Delta G^\circ = 0$	zero	$K = 1$	Neither side favoured; balanced
$\Delta G^\circ > 0$ (positive)	negative	$K < 1$	Reactants favoured; forward reaction not spontaneous

The more negative ΔG°, the larger K, and the further towards products the equilibrium lies. Qualitatively, this explains why some reactions go essentially to completion (very negative ΔG°) while others barely

proceed.

How K varies with temperature

Because $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$, the temperature dependence of K depends on the sign of ΔH°. Combining the two relationships gives the qualitative **van't Hoff** result: for an **endothermic** reaction (ΔH° positive) K **increases** as temperature rises, whereas for an **exothermic** reaction (ΔH° negative) K **decreases** as temperature rises. This agrees exactly with Le Châtelier's prediction in Section 4 and confirms that temperature is the only variable that changes K.

K_c and K_p — a note

For gas equilibria, K_p (built from partial pressures) and K_c (built from concentrations) describe the same equilibrium and are related through the ideal gas law; they are equal in value only when the number of moles of gas is the same on both sides ($\Delta n = 0$). Both increase or decrease with temperature in the same qualitative way, governed by the sign of ΔH°.

Worked example 8 (HL) — ΔG° and K linked

A reaction has $\Delta G^\circ = -11.4$ kJ mol^{-1} at 298 K. Calculate K.

Rearrange: $\ln K = -\Delta G^\circ / (RT) = -(-11400) / (8.31 \times 298)$.

$\ln K = 11400 / 2476 = 4.60$, so $K = e^{4.60} =$ **100** (2 s.f.).

$K > 1$ and ΔG° is negative — consistent: the products are favoured.

9. Common pitfalls

- **Including pure solids or liquids** in a K expression. Only gases and aqueous species appear.
- **Forgetting that coefficients become powers**. [HI] appears squared because of the 2 in 2HI.
- **Claiming a concentration or pressure change alters K**. Only temperature changes K.
- **Confusing Q with K**. Q uses the current concentrations; the comparison Q vs K predicts the shift.
- **Saying a catalyst increases yield**. It only speeds up attainment of the same equilibrium.
- **Applying pressure effects when moles of gas are equal**, or resolving pressure onto species in the condensed phase.
- **Confusing extent with rate**. A large K says a reaction goes far, not that it goes fast.

10. Quick reference

Idea	Statement
Equilibrium	Closed system; forward rate = reverse rate; macroscopic properties constant; dynamic
K_c expression	$K_c = [\text{products}]^{\text{coeff}} / [\text{reactants}]^{\text{coeff}}$; omit pure solids and liquids
Magnitude of K	$K > 1$ favours products; $K < 1$ favours reactants; $K \approx 1$ similar amounts
Reaction quotient	$Q < K \rightarrow$ shift forward; $Q > K \rightarrow$ shift back; $Q = K \rightarrow$ at equilibrium
Le Châtelier	System opposes an imposed change (concentration, pressure, temperature)

Idea	Statement
What changes K	Temperature only; concentration, pressure and catalysts do not
Catalyst	No change to position or K; equilibrium reached faster
Industry	Conditions are a compromise between yield, rate and cost
(HL) Gibbs link	$\Delta G^\circ = -RT \ln K$; more negative ΔG° → larger K

11. Test yourself

Attempt all ten without notes, then check against the full answers that follow.

1. Write the K_c expression for $2NO_2(g) \rightleftharpoons N_2O_4(g)$.

2. Write the K_c expression for $Fe_2O_3(s) + 3CO(g) \rightleftharpoons 2Fe(s) + 3CO_2(g)$.

3. State and explain the effect on the position of equilibrium of increasing the pressure on $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$.

4. For an endothermic reaction, state and explain the effect of raising the temperature on the yield and on K.

5. A reaction has $K_c = 3.5 \times 10^{-8}$. Comment on the extent of reaction.

6. In the Haber process, explain why a temperature of about 450 °C is used rather than a much lower temperature.

7. At equilibrium a 2.0 dm^3 vessel holds 0.40 mol PCl_5, 0.20 mol PCl_3 and 0.20 mol Cl_2 for $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$. Find K_c.

8. 1.00 mol of N_2O_4 is placed in a 1.00 dm^3 flask and 0.60 mol dissociates: $N_2O_4(g) \rightleftharpoons 2NO_2(g)$. Use an ICE table to find K_c.

9. For $H_2 + I_2 \rightleftharpoons 2HI$, $K_c = 49$. Equal amounts, 0.300 mol each of H_2 and I_2, are mixed in a 1.00 dm^3 flask. Find [HI] at equilibrium.

10. (HL) A reaction has $\Delta G^\circ = +5.7$ kJ mol^{-1} at 298 K. Calculate K and comment.

Answers

1. $K_c = [N_2O_4] / [NO_2]^2$. The coefficient 2 becomes the power on NO_2.

2. $K_c = [CO_2]^3 / [CO]^3$. The solids Fe_2O_3 and Fe are omitted.

3. The equilibrium shifts **forward** (to the right). There are 4 mol of gas on the left and 2 on the right, so the system moves to the side with fewer moles of gas to oppose the pressure increase; more NH_3 forms. K is unchanged.

4. Raising the temperature shifts an endothermic reaction **forward** (heat behaves like a reactant), so the yield **increases** and K **increases**. Temperature is the only factor that changes K.

5. K_c is extremely small (<< 1), so the equilibrium lies far to the left; only a tiny amount of product forms — the reaction barely proceeds.

6. A lower temperature would give a higher equilibrium yield (the reaction is exothermic) but the rate would be impractically slow. 450 °C is a **compromise** that gives an acceptable rate at a still-usable yield.

7. Concentrations: $[PCl_5] = 0.40/2.0 = 0.20$; $[PCl_3] = [Cl_2] = 0.20/2.0 = 0.10$ mol dm^{-3}. $K_c = (0.10 \times 0.10) / 0.20 =$ **0.050 mol dm^{-3}**.

8. ICE: N_2O_4 1.00 → 0.40; NO_2 0 → 2(0.60) = 1.20. $K_c = (1.20)^2 / 0.40 = 1.44 / 0.40 =$ **3.6 mol dm^{-3}**.

9. Let x react: $[HI] = 2x$, $[H_2] = [I_2] = 0.300 - x$. $\sqrt{49} = 7 = 2x/(0.300 - x)$, so $2x = 2.10 - 7x$, $9x = 2.10$, $x = 0.2333$. $[HI] = 2x =$ **0.467 mol dm^{-3}**.

10. $\ln K = -\Delta G°/RT = -5700 / (8.31 \times 298) = -2.30$, so $K = e^{-2.30} =$ **0.10**. $K < 1$ and $\Delta G°$ is positive: reactants are favoured, the forward reaction is not spontaneous under standard conditions.