

Equilibria

AS

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Last checked 30 September 2026

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What you need to know

By the end of this topic you should be able to:

7.1 Chemical equilibria

1. **Explain what a reversible reaction is:** one that can go both ways, so products can turn back into reactants.
2. **Explain dynamic equilibrium:** the forward and reverse reactions keep happening at the same rate, so the concentrations of reactants and products stay constant.
3. **Explain why a closed system is needed** for dynamic equilibrium to be reached.
4. **State Le Chatelier's principle:** if a change is made to a system at dynamic equilibrium, the position of equilibrium moves to minimise that change.
5. **Use Le Chatelier's principle** to predict, and explain, what a change in temperature, concentration or pressure, or adding a catalyst, does to a system at equilibrium.
6. **Write the expression for K_c ,** the equilibrium constant in terms of concentrations, from an equation, and work out its units.
7. **Use the terms mole fraction and partial pressure.**
8. **Write the expression for K_p ,** the equilibrium constant in terms of partial pressures, and work out its units. (You do not need the link between K_p and K_c .)
9. **Calculate with K_c and K_p :** a value of K from equilibrium data, or an unknown amount from K . No quadratic equations are needed.
10. **Calculate the amounts present at equilibrium** from starting amounts and one equilibrium amount.
11. **State whether a change of temperature, concentration or pressure, or a catalyst, changes the value of K .**
12. **Describe and explain the conditions used in the Haber process and the Contact process,** as examples of Le Chatelier's principle at work in industry.

7.2 Brønsted–Lowry acids and bases

13. **Name and write the formulas of the common acids:** hydrochloric acid HCl, sulfuric acid H₂SO₄, nitric acid HNO₃, ethanoic acid CH₃COOH.

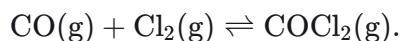
14. **Name and write the formulas of the common alkalis:** sodium hydroxide NaOH, potassium hydroxide KOH, ammonia NH₃.
15. **Describe the Brønsted–Lowry theory:** an acid is a proton donor and a base is a proton acceptor.
16. **Describe strong acids and bases** as fully dissociated in water, and **weak acids and bases** as only partly dissociated.
17. **Know the pH scale:** water has pH 7, acids below 7, alkalis above 7.
18. **Explain how strong and weak acids behave differently:** with a reactive metal, and in pH (pH meter or universal indicator) and conductivity.
19. **Understand neutralisation** as H⁺(aq) and OH[−](aq) forming H₂O(l), and that a salt forms too.
20. **Sketch pH titration curves** for every pairing of strong or weak acid with strong or weak alkali.
21. **Choose a suitable indicator** for a titration from data about indicator colour-change ranges (p*K*_a values are not used).

Calculating pH from concentrations, *K*_a, buffers and *K*_{sp} are **not** in this topic: they come in the A Level topic on equilibria.

Understand it

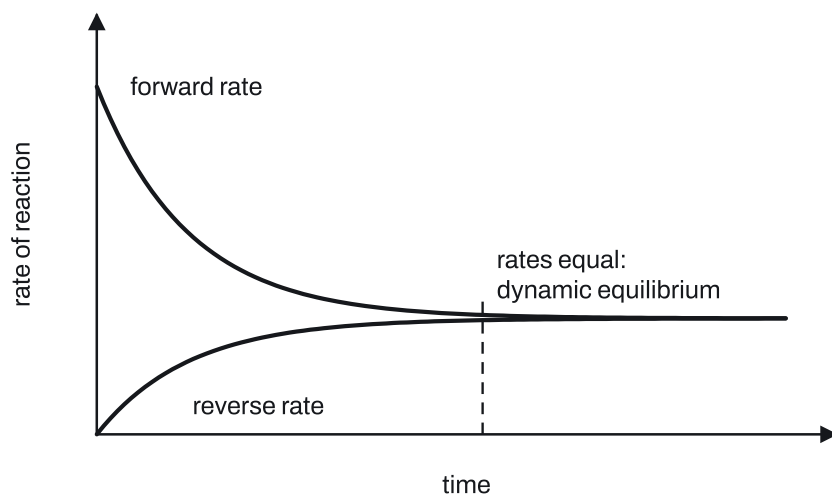
Reversible reactions and dynamic equilibrium

Many reactions can go both ways. We write them with the sign $\rightleftharpoons$, for example



The reaction left to right is the **forward reaction**; right to left is the **reverse** (or backward) reaction.

Start with only reactants in a sealed flask. At first the forward reaction is fast, because the reactants are concentrated, and the reverse reaction has a rate of zero, because there is no product yet. As reactants are used up the forward rate falls; as product builds up the reverse rate rises. Eventually the two rates become **equal**. From then on, every second, as many molecules are made as are broken back down, so the **concentrations stay constant**, even though both reactions are still going on. This is **dynamic equilibrium**: "dynamic" because the reactions never stop.



Equilibrium needs a **closed system**: nothing can get in or out. If a gaseous product escapes from an open tube, the reverse reaction can't happen at the right rate, so the reaction just runs one way until a reactant runs out. A reaction can be reversible without ever reaching equilibrium.

Le Chatelier's principle

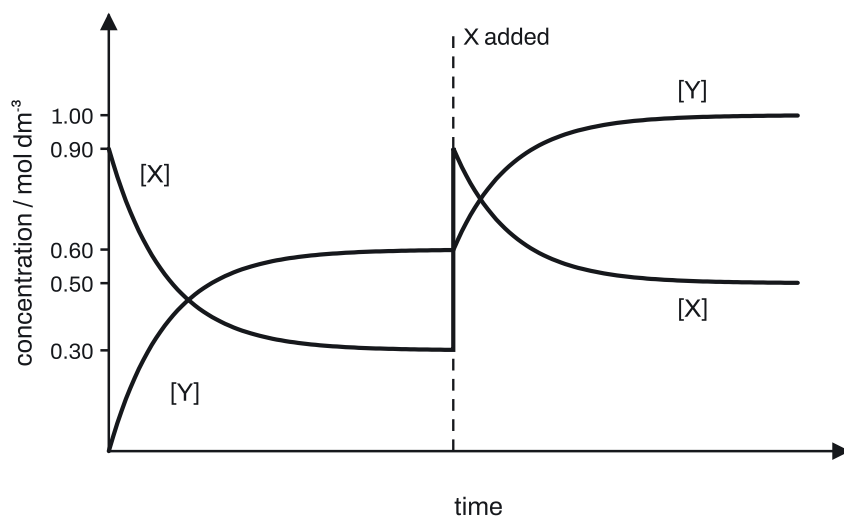
The **position of equilibrium** describes how much of the mixture is products and how much is reactants. "Moves to the right" means more products form; "moves to the left" means more reactants.

Le Chatelier's principle: if a change is made to a system at dynamic equilibrium, the position of equilibrium moves to minimise that change. The system partly "undoes" what you did to it. It never fully undoes it.

Change	What the equilibrium does
Add a reactant (or remove a product)	Moves right to use up some of the added reactant
Add a product (or remove a reactant)	Moves left
Raise the temperature	Moves in the endothermic direction, which takes in heat
Lower the temperature	Moves in the exothermic direction, which gives out heat
Raise the pressure (gases)	Moves to the side with fewer moles of gas
Lower the pressure	Moves to the side with more moles of gas
Add a catalyst	No change in position; equilibrium is reached faster

A catalyst speeds up the forward and reverse reactions by the same factor, so it can't favour either side. If both sides have the same number of moles of gas, pressure has no effect on the position. Only count **gases** when you compare sides: solids and liquids hardly change volume with pressure. Adding an unreactive gas at constant volume changes no concentrations, so the position doesn't move.

The sign of ΔH always refers to the forward reaction. If ΔH is negative, the forward reaction is exothermic and the reverse is endothermic, so raising the temperature moves the position to the left and lowers the yield.



The graph shows the principle with numbers. For $X \rightleftharpoons Y$, where the ratio $[Y] \div [X]$ is always 2 at equilibrium, adding 0.60 mol dm^{-3} of X raises $[X]$ from 0.30 to 0.90. The equilibrium moves right and uses some of it up, but $[X]$ settles at 0.50: lower than just after the addition, higher than before it. The change is minimised, not reversed.

Coloured equilibria make these shifts visible. For



warming turns the mixture more blue (the endothermic forward reaction is favoured), cooling turns it more pink, and adding chloride ions (concentrated hydrochloric acid) turns it more blue.

The equilibrium constant K_c

For a general reaction $aA + bB \rightleftharpoons cC + dD$ at equilibrium,

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

where $[]$ means the concentration at equilibrium in mol dm^{-3} . **Products go on top, reactants underneath, and each is raised to the power of its number in the equation.** The terms are multiplied, never added.

For the reaction above, $K_c = \frac{[\text{CoCl}_2]}{[\text{CO}][\text{Cl}_2]}$.

At a fixed temperature K_c has the same value for every equilibrium mixture of that reaction, whatever amounts you started with. A large K_c (much more than 1) means the equilibrium lies to the right; a small one means it lies to the left.

Leave out solids, and leave out water when it is the solvent in an aqueous reaction: their concentrations don't change. When every substance is a liquid mixed together (as when an alcohol and a carboxylic acid make an ester and water), water **is** included.

Units. Put mol dm^{-3} in for every concentration and cancel.

For $2X(g) + Y(g) \rightleftharpoons 2Z(g)$: $K_c = \frac{[Z]^2}{[X]^2[Y]}$, units = $\frac{(\text{mol dm}^{-3})^2}{(\text{mol dm}^{-3})^3} = (\text{mol dm}^{-3})^{-1} = \text{dm}^3 \text{mol}^{-1}$.

A quick way: units = $(\text{mol dm}^{-3})^n$ where n = (sum of the powers on top) – (sum of the powers underneath). If $n = 0$, K_c has **no units**.

Mole fraction, partial pressure and K_p

In a mixture of gases, each gas pushes on the walls as if it were there alone. The pressure it would exert is its **partial pressure**, and the partial pressures add up to the total pressure.

$$\text{mole fraction of A, } x_A = \frac{\text{moles of A}}{\text{total moles of gas}} \quad p_A = x_A \times p_{\text{total}}$$

The mole fractions of all the gases add up to 1. For gas reactions, the equilibrium constant can use partial pressures instead of concentrations:

$$K_p = \frac{p_C^c p_D^d}{p_A^a p_B^b}$$

Only **gases** appear in K_p . Its units are worked out the same way, with Pa, kPa or atm in place of mol dm^{-3} : for $2X(g) + Y(g) \rightleftharpoons 2Z(g)$, K_p has units of Pa^{-1} (or kPa^{-1} , atm^{-1}).

What changes the value of K ?

Only temperature. Changing concentration or pressure moves the position of equilibrium, but the mixture settles with the same value of K . A catalyst changes neither the position nor K .

Forward reaction	Raise the temperature	Lower the temperature
Exothermic ($\Delta H < 0$)	K decreases	K increases
Endothermic ($\Delta H > 0$)	K increases	K decreases

This works backwards too: if K gets bigger when the temperature rises, the forward reaction is endothermic.

The Haber and Contact processes

Both are exothermic reactions with fewer moles of gas on the right, so Le Chatelier says **low temperature and high pressure** give the best yield. In practice a compromise is used.

	Haber process	Contact process
Reaction	$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, $\Delta H = -92 \text{ kJ mol}^{-1}$	$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$, $\Delta H = -197 \text{ kJ mol}^{-1}$
Mole ratio of reactants	$N_2 : H_2 = 1 : 3$	$SO_2 : O_2 = 2 : 1$ (excess air is used)
Temperature	about 400–450 °C	about 400–450 °C

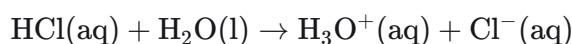
	Haber process	Contact process
Pressure	about 200 atm (20 MPa)	about 1–2 atm (100–200 kPa)
Catalyst	iron, Fe (a solid: heterogeneous)	vanadium(V) oxide, V_2O_5 (a solid: heterogeneous)

Why these conditions:

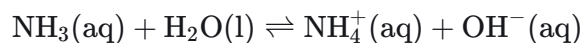
- **Temperature.** A lower temperature gives a higher equilibrium yield, but the reaction would be far too slow. About 450 °C is a compromise: a reasonable yield made quickly. The catalyst also only works well when hot.
- **Pressure (Haber).** High pressure moves the position right (4 moles of gas become 2) and speeds up the reaction, but very high pressure is expensive to generate and needs thick, strong equipment. About 200 atm is a compromise.
- **Pressure (Contact).** The yield is already very high (well over 90%) at 1–2 atm, so extra pressure would cost a lot for almost no gain.
- **Excess oxygen (Contact).** Adding more of a cheap reactant moves the position right and uses up more of the valuable SO_2 .
- **Catalyst.** It makes equilibrium arrive faster, so more product is made per hour, but it does not change the yield at equilibrium or the value of K . Unreacted gases are recycled.

Brønsted–Lowry acids and bases

A **Brønsted–Lowry acid** is a proton donor; a **Brønsted–Lowry base** is a proton acceptor. A "proton" here is an H^+ ion. When hydrogen chloride dissolves in water, it donates a proton to water:



HCl is the acid and water is the base. Ammonia accepts a proton from water, so here water is the acid:



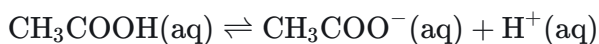
Water can act as either, depending on what it meets. When an acid loses its proton, what is left can take a proton back, so it is a base: the **conjugate base** of that acid (Cl^- from HCl, CH_3COO^- from CH_3COOH). Likewise NH_4^+ is the **conjugate acid** of NH_3 . An acid and its conjugate base differ by exactly one H^+ .

Any species that takes a proton is acting as a base, so metal oxides are bases too: an oxide ion accepts protons from an acid to form water.

Acids	Formula	Strength	Alkalis	Formula	Strength
hydrochloric acid	HCl	strong	sodium hydroxide	NaOH	strong
sulfuric acid	H_2SO_4	strong	potassium hydroxide	KOH	strong
nitric acid	HNO_3	strong	ammonia	NH_3	weak
ethanoic acid	CH_3COOH	weak			

Strong and weak

Strong acids and bases are **fully dissociated** (split into ions) in water: every HCl gives an H^+ . **Weak** acids and bases are only **partly dissociated**: most molecules stay whole, and the ions are in equilibrium with them.



Strength (how much it dissociates) is not the same as concentration (how much is dissolved per dm^3). A dilute strong acid and a concentrated weak acid can have the same pH.

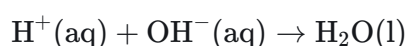
Compare $0.100 \text{ mol dm}^{-3}$ hydrochloric acid with $0.100 \text{ mol dm}^{-3}$ ethanoic acid:

Test	Strong acid (HCl)	Weak acid (CH_3COOH)	Why
pH meter	pH 1.0	pH about 3	fewer H^+ ions in the weak acid
universal indicator	red	orange	same reason
conductivity	high	low	fewer ions to carry the current
reaction with magnesium	fast fizzing	slower fizzing	lower $[\text{H}^+]$, so fewer collisions per second
total H_2 made (same volume, excess Mg)	same	same	as H^+ is used up, the equilibrium moves right and releases more, until all the acid has reacted

The pH scale: pure water has pH 7 (neutral), acidic solutions have pH below 7, alkaline ones above 7. The lower the pH, the higher $[\text{H}^+]$.

Neutralisation

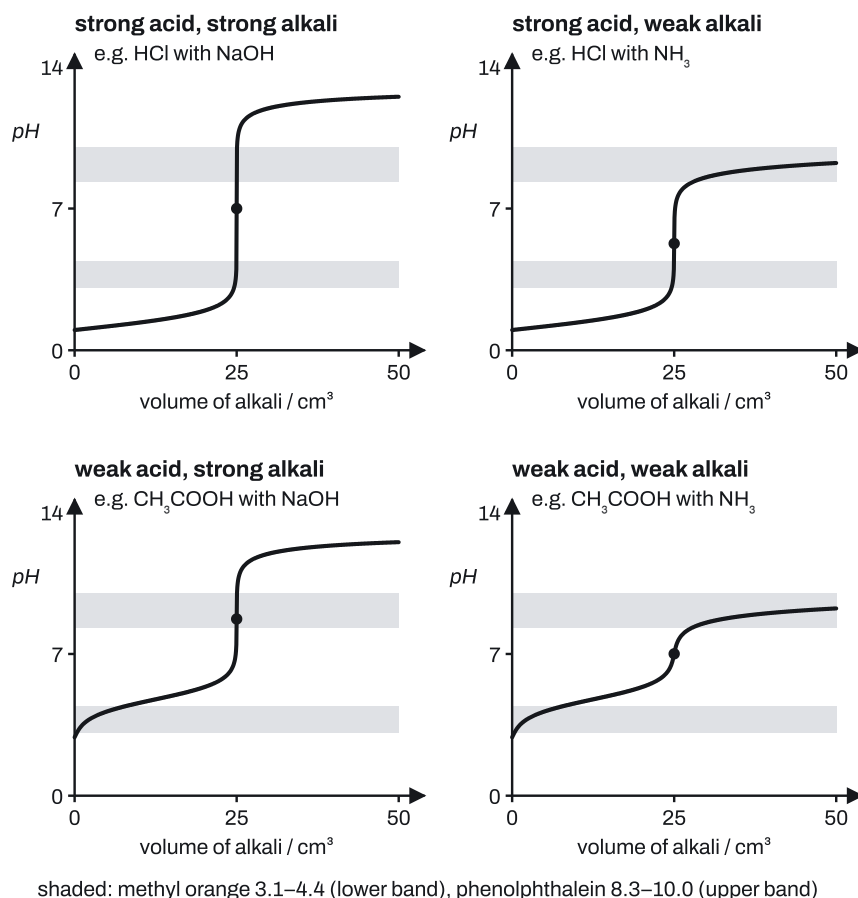
In every acid–alkali reaction the real change is



The other ions make a **salt**: $\text{HNO}_3 + \text{KOH} \rightarrow \text{KNO}_3 + \text{H}_2\text{O}$. With ammonia the salt forms directly: $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4\text{Cl}$. Watch the ratios: H_2SO_4 gives two H^+ , so it needs two NaOH.

pH titration curves

A pH titration curve plots pH against the volume of alkali added (or acid, if the acid is added from the burette). With equal concentrations and 25.0 cm^3 of acid, the **equivalence point** is at 25.0 cm^3 . Near it, a tiny volume changes the pH a lot, so the curve has a near-vertical section.



Read each curve from three features: where it **starts** (a strong acid starts near pH 1, a weak acid higher, near 3), where the **vertical section** is (its middle is the pH at equivalence), and where it **ends** (a strong alkali levels off near 12–13, a weak one near 9–10).

Acid + alkali	Vertical section (about)	pH at equivalence
strong + strong	3 to 11	7
strong + weak	3 to 7	below 7
weak + strong	7 to 11	above 7
weak + weak	none	about 7

Choosing an indicator

An indicator changes colour over a small range of pH. It gives a sharp end-point only if its **whole range lies on the vertical section** of the curve, because then one drop at the equivalence point takes the pH right through the range.

- strong acid + strong alkali: methyl orange (3.1–4.4) or phenolphthalein (8.3–10.0) both work;
- strong acid + weak alkali: methyl orange;
- weak acid + strong alkali: phenolphthalein;
- weak acid + weak alkali: no indicator is suitable.

In the exam you are given each indicator's range and choose the one that fits.

Key facts and formulas

Nothing in this topic is printed in the 9701 data booklet, so learn every line.

Formula	Status
For $aA + bB \rightleftharpoons cC + dD$: $K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$ (leave out solids)	Learn it
$K_p = \frac{p_C^c p_D^d}{p_A^a p_B^b}$ (gases only)	Learn it
Units of K_c : $(\text{mol dm}^{-3})^{\Delta n}$; of K_p : $(\text{Pa})^{\Delta n}$, with Δn = powers on top – powers underneath	Learn it
Concentration = $\frac{\text{moles}}{\text{volume in dm}^3}$	Learn it
Mole fraction $x_A = \frac{n_A}{n_{\text{total}}}$; the mole fractions add up to 1	Learn it
Partial pressure $p_A = x_A \times p_{\text{total}}$; the partial pressures add up to p_{total}	Learn it
Only temperature changes K ; raising T increases K for an endothermic forward reaction and decreases it for an exothermic one	Learn it
Neutralisation: $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$	Learn it
Acid = proton donor; base = proton acceptor; strong = fully dissociated, weak = partly dissociated	Learn it
Haber: Fe, about 450 °C, about 200 atm; Contact: V_2O_5 , about 450 °C, 1–2 atm	Learn it

How to do it

We have tagged 96 questions on this topic from Papers 1 and 2, 2021–2025. 21 of them mix more than one type, so the counts add up to more than 96. Types 1–5 and 7–9 are syllabus 7.1 (equilibria); types 6 and 10 are 7.2 (acids and bases).

Question type	Questions
Predict and explain the effect of a change on position, yield or colour (7.1)	37
Calculate equilibrium amounts and K_c from starting amounts (7.1)	18

Question type	Questions
Write a K_c or K_p expression and its units (7.1)	15
Mole fractions, partial pressures and K_p (7.1)	14
The effect of a change on the value of K (7.1)	12
Brønsted–Lowry acids and bases, strong and weak (7.2)	12
Define dynamic equilibrium, closed system, Le Chatelier's principle (7.1)	11
Haber and Contact process conditions (7.1)	6
Use a given K to find an unknown amount (7.1)	5
Sketch or identify pH titration curves and indicators (7.2)	5

1. Predict and explain the effect of a change on position, yield or colour (37 questions)

- Find the facts in the equation:** the sign of ΔH (for the forward reaction), the number of moles of **gas** on each side, and which species is coloured.
- Identify the change** and what would "minimise" it: extra heat is absorbed by the endothermic direction; extra pressure is reduced by making fewer gas molecules; an added substance is used up.
- State the direction** the position moves (left or right).
- Give the consequence** asked for: more or less product (yield), or the colour becoming "more X".
- Explain with a reason tied to the equation**, such as "the forward reaction is endothermic" or "there are fewer moles of gas on the right".

On Paper 2 each change usually earns two marks: one for the observation (the colour) and one for the reason. The observation is about the **mixture**, so write "the mixture becomes more blue", not "blue product forms".

Variations you will meet:

- Best conditions for yield** (Paper 1): pick high or low pressure from the gas moles, and high or low temperature from the sign of ΔH . A catalyst never increases the yield.
- Deduce ΔH and the gas moles from a graph** of yield against pressure at two temperatures. If yield rises with pressure, the product side has fewer moles of gas. If the lower temperature gives the higher yield, the forward reaction is exothermic.
- Pick the right graph:** a catalyst gives the same final amount, sooner; a lower temperature (for an exothermic reaction) gives a larger final amount, reached more slowly.
- New concentration after adding a substance.** The position moves to use up only **part** of what was added, so the new equilibrium concentration is **above** the old equilibrium value but **below** the old value plus (moles added $\div$ volume). For example, adding 0.40 mol of W to 2.0 dm³ of a mixture with $[W] = 0.50$ mol dm⁻³ raises $[W]$ at first by $0.40 \div 2.0 = 0.20$, so the new equilibrium $[W]$ lies between 0.50 and 0.70 mol dm⁻³.

- **New concentrations after a temperature change**, when one new value is given. Use the mole ratios. For $A + B \rightleftharpoons 2C$ with equilibrium concentrations 0.50, 0.30 and 1.20 mol dm⁻³, if heating makes $[C]$ fall to 1.00, then 0.20 of C has become 0.10 of A and 0.10 of B: $[A] = 0.60$ and $[B] = 0.40$ mol dm⁻³.
- **An unreactive gas added at constant volume**: no change, because no concentration or partial pressure changes.
- **Adding an ion already in a weak-acid equilibrium**. For $HA \rightleftharpoons H^+ + A^-$, adding the salt NaA raises $[A^-]$, so the position moves left, $[H^+]$ falls and the pH rises. Removing OH^- or H^+ (for example by adding an acid that reacts with it) moves an equilibrium towards the side that makes more of it.

2. Calculate equilibrium amounts and K_c from starting amounts (18 questions)

1. **Write three rows** under the equation: initial moles, change, equilibrium moles.
2. **Fill in the change from the one equilibrium amount you know**, using the mole ratios in the equation. Reactants go down, products go up.
3. **Divide each equilibrium amount by the volume** (in dm³) to get concentrations.
4. **Substitute into K_c** , give the value to 3 significant figures and add the units.

For example, $P(g) + 2Q(g) \rightleftharpoons R(g)$ in a 2.0 dm³ flask, starting with 0.60 mol of P and 0.40 mol of Q; at equilibrium there is 0.10 mol of R.

	P	Q	R
initial / mol	0.60	0.40	0
change / mol	-0.10	-0.20	+0.10
equilibrium / mol	0.50	0.20	0.10
concentration / mol dm ⁻³	0.25	0.10	0.050

$$K_c = \frac{0.050}{0.25 \times 0.10^2} = 20 \text{ dm}^6 \text{ mol}^{-2}.$$

Variations you will meet:

- **Some product present at the start**: it goes in the "initial" row; the change row is worked out exactly the same way.
- **Two moles of product per mole of reactant**: if 0.05 mol of H₂ reacts in $H_2 + Br_2 \rightleftharpoons 2HBr$, 0.10 mol of HBr forms.
- **Equal moles on both sides** (the powers on top and underneath add up to the same number): the volumes cancel, so you can put moles straight into K_c . If they are not equal, you must divide by the volume.
- **An amount found from other data**: a titration, or a mass change divided by M_r . Work out moles first, then use the table.
- **Working backwards** to a starting amount: put the unknown in the initial row, write the equilibrium amounts in terms of it, and solve.
- **An expression in terms of a letter**: if the mole fraction of product is x , write each amount in terms of x and substitute. Paper 1 usually asks you to pick the matching expression.

3. Write a K_c or K_p expression and its units (15 questions)

1. **Products over reactants**, each raised to the power of its coefficient; multiply the terms.
2. Use **square brackets** for K_c and p for K_p . Leave out solids; for K_p include only gases.
3. **Units**: count the powers on top minus the powers underneath, Δn . K_c has units $(\text{mol dm}^{-3})^{\Delta n}$ and K_p has units $(\text{Pa})^{\Delta n}$ (or kPa, atm). If $\Delta n = 0$ there are no units.

Examples: for $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$, $\Delta n = 2 - 4 = -2$, so K_c is in $\text{mol}^{-2} \text{dm}^6$ and K_p in Pa^{-2} . For $\text{H}_2 + \text{Br}_2 \rightleftharpoons 2\text{HBr}$, $\Delta n = 0$: no units.

Variations: "Which reaction's K has no units?" means find the one with equal gas moles on both sides. A mark scheme accepts the units written either way, such as $\text{mol}^{-2} \text{dm}^6$ or $\text{dm}^6 \text{mol}^{-2}$.

4. Mole fractions, partial pressures and K_p (14 questions)

1. **Find the equilibrium moles of every gas** (a moles table, as in type 2, if you're given starting amounts).
2. **Add them up** to get the total moles of gas.
3. **Mole fraction** of each gas = its moles $\div$ total.
4. **Partial pressure** = mole fraction $\times$ total pressure.
5. **Substitute into K_p** and give units.

For instance, a gas mixture of 2.0 mol at a total pressure of 200 kPa that contains 0.50 mol of X: $x_X = 0.25$ and $p_X = 50 \text{ kPa}$.

Variations you will meet:

- **Answers in terms of p** (Paper 1): work out the moles, find the mole fraction as a fraction, and multiply by p .
- **Missing partial pressure**: subtract the known partial pressures from the total, or rearrange K_p .
- **From a partial pressure back to moles**: moles of A = $\frac{p_A}{p_{\text{total}}} \times \text{total moles}$.
- **A decomposition** such as $\text{X(g)} \rightleftharpoons \text{Y(g)} + \text{Z(g)}$: Y and Z always have equal partial pressures, which saves work.
- **Checking statements** about an equilibrium mixture: the gas with more moles has the larger partial pressure, and no mole fraction can be more than 1.

5. The effect of a change on the value of K (12 questions)

1. **Is the change a temperature change?** If not (pressure, concentration, catalyst), the value of K does **not** change. Say so and give the reason: " K depends only on temperature".
2. **If it is temperature**, use the sign of ΔH : raising T moves the position in the endothermic direction. If that is the forward direction, K increases; if it's the reverse, K decreases.

Variations:

- **Work back from a change in K :** K rose, so either the temperature rose and the forward reaction is endothermic, or the temperature fell and it is exothermic. In a multiple-choice question, a change of K rules out every pressure or concentration option.
- **Compare K at two temperatures** to find the sign of ΔH : if K is larger at the higher temperature, the forward reaction is endothermic.
- **"Moves the position right and increases K ":** only a temperature change can do both.

6. Brønsted–Lowry acids and bases, strong and weak (12 questions)

1. **Definitions** (1 mark each): an acid is a **proton donor**, a base is a **proton acceptor**. A strong acid is a proton donor that is **fully dissociated** in water; a weak acid is **partly dissociated**.
2. **Which species is the acid or base in an equation?** Follow the H^+ : the species that loses it is the acid, the one that gains it is the base.
3. **Conjugate pairs:** remove one H^+ from the acid (lower the charge by 1) to get its conjugate base: $H_2SO_4 \rightarrow HSO_4^-$; add one to a base to get its conjugate acid: $NH_3 \rightarrow NH_4^+$.
4. **Equations for weak acids and bases with water** use $\rightleftharpoons$: $HA + H_2O \rightleftharpoons H_3O^+ + A^-$, and for a base $B + H_2O \rightleftharpoons BH^+ + OH^-$ (an anion base A^- gives $HA + OH^-$).

Variations you will meet:

- **Explain pH values** of solutions of the same concentration: strong acid lowest pH (full dissociation, most H^+), weak acid higher; strong alkali highest pH, weak alkali lower.
- **Strong vs weak with a reactive metal:** the weak acid reacts more slowly (lower $[H^+]$) but, with the same moles of acid, gives the same total volume of gas.
- **Water as an acid or a base:** it is a base when it gains H^+ (becoming H_3O^+) and an acid when it loses one (becoming OH^-).
- **Why an ammonium salt dissolves a basic solid:** NH_4^+ donates a proton to the OH^- , removing it, so the solid's equilibrium moves to replace it.
- **Oxides as bases:** a metal oxide accepts protons from an acid.
- **Misleading statements** in Paper 1: the number of H atoms in a formula does not decide strength, and "concentrated" does not mean "strong".

7. Define dynamic equilibrium, closed system, Le Chatelier's principle (11 questions)

1. **Dynamic equilibrium** (1 mark): the rates of the forward and reverse reactions are equal, **or** the concentrations of reactants and products stay constant. Writing both is safest.
2. **The condition needed** (1 mark): a **closed system**.
3. **Le Chatelier's principle** (2 marks): if a change is made to a system at dynamic equilibrium, **the position of equilibrium moves (1) to minimise the change (1)**.

Variations:

- **Sketch the reverse rate** on a rate–time graph: it starts at **zero**, rises with a decreasing slope, and meets the falling forward rate at the same level; after that both lines are flat and together (see the first diagram).
- **Colour at the start compared with at equilibrium**: at equilibrium both coloured species are present, so the colour is in between and then stops changing.
- **Is a reaction at equilibrium?** A solid that decomposes in an open tube, with a product escaping or re-forming somewhere else, is reversible but not at equilibrium, because the system is open.

8. Haber and Contact process conditions (6 questions)

1. **Recall the table** in "Understand it": catalyst, temperature, pressure, ratio of reactants.
2. **Explain each condition** as a balance between yield (Le Chatelier), rate and cost.

Variations: statements to judge in Paper 1: iron is a **heterogeneous** catalyst (a solid with gases); the catalyst leaves the position and K_p unchanged; K_p doesn't change with pressure; Haber K_p has units (Pa^{-2}); the Contact yield is already high at low pressure; raising the temperature lowers the yield of both. In a plant diagram, a heat exchanger warms the incoming gases using the heat given out by the reaction, and a condenser removes ammonia as a liquid so unreacted gases can be recycled.

9. Use a given K to find an unknown amount (5 questions)

1. **Write the K expression** with the numbers you know.
2. **Rearrange** for the unknown. If the unknown appears squared (or two unknowns are equal, so their product is a square), take the square root. No quadratics are needed.
3. **Convert** to the unit asked: concentration $\times$ volume gives moles.

For $\text{A} + \text{B} \rightleftharpoons 2\text{D}$ with $K_c = 16.0$, $[\text{A}] = 0.25$ and $[\text{B}] = 0.64 \text{ mol dm}^{-3}$: $[\text{D}]^2 = 16.0 \times 0.25 \times 0.64 = 2.56$, so $[\text{D}] = 1.60 \text{ mol dm}^{-3}$.

Variations: a ratio of two concentrations is given (substitute one in terms of the other); equal starting amounts of two reactants with 1 : 1 stoichiometry stay equal at equilibrium; an unknown starting amount is found by working back through a moles table.

10. Sketch or identify pH titration curves and indicators (5 questions)

1. **Decide the strength** of the acid and the alkali.
2. **Starting pH**: about 1 for a strong acid, about 3 for a weak acid (both 0.1 mol dm^{-3}).
3. **Equivalence volume** from the moles and the ratio in the equation.
4. **Vertical section at that volume**: about 3–11 (strong + strong), 3–7 (strong acid + weak alkali), 7–11 (weak acid + strong alkali), none (weak + weak).
5. **Final pH**, well after the end point: about 12–13 for a strong alkali, 9–10 for a weak one.
6. **Indicator**: the one whose colour-change range lies wholly within the vertical section.

Variations:

- **Identify the pair from a curve** (Paper 1): read the starting pH, the vertical section and the final pH, then check the equivalence volume. An alkali with two OH^- per formula, such as $\text{Ba}(\text{OH})_2$, at the same concentration as HCl needs only **half** the volume of acid; a diprotic acid like H_2SO_4 needs **twice** the volume of NaOH .
- **Acid added to alkali**: the curve is flipped, starting high and falling.

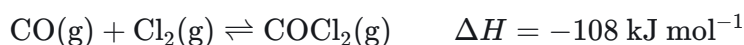
Worked examples

The bold codes show where marks are earned, as in Cambridge mark schemes. The number is how many marks.

- **M1** method mark: for a correct method, even if the numbers slip.
 - **A1** accuracy mark: for a correct answer; it needs the M mark before it.
 - **B1** mark for a correct result on its own, with no method needed.
-

Example 1

Carbonyl chloride, COCl_2 , is made from carbon monoxide and chlorine.



0.500 mol of CO and 0.400 mol of Cl_2 are sealed in a 2.00 dm^3 flask and left to reach equilibrium at a constant temperature. The equilibrium mixture contains 0.150 mol of COCl_2 .

- Write the expression for K_c and give its units. **[2]**
- Calculate the value of K_c at this temperature. Give your answer to three significant figures. **[3]**
- State the effect, if any, on the value of K_c of increasing the pressure at constant temperature. **[1]**
- The temperature is raised. State and explain the effect on the amount of COCl_2 at equilibrium and on the value of K_c . **[2]**

(a) $K_c = \frac{[\text{COCl}_2]}{[\text{CO}][\text{Cl}_2]}$ **B1**. Units: $\frac{\text{mol dm}^{-3}}{(\text{mol dm}^{-3})^2} = \text{dm}^3 \text{ mol}^{-1}$ **B1**.

(b) 0.150 mol of COCl_2 formed, so 0.150 mol of each reactant was used: $\text{CO} = 0.500 - 0.150 = 0.350 \text{ mol}$ and $\text{Cl}_2 = 0.400 - 0.150 = 0.250 \text{ mol}$. **M1**

Concentrations in 2.00 dm^3 : $[\text{CO}] = 0.175$, $[\text{Cl}_2] = 0.125$, $[\text{COCl}_2] = 0.0750 \text{ mol dm}^{-3}$. **M1**

$$K_c = \frac{0.0750}{0.175 \times 0.125} = 3.43 \text{ dm}^3 \text{ mol}^{-1}$$

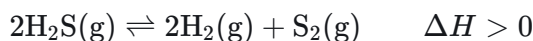
A1

(c) No effect: K_c only changes with temperature. **B1** (The position moves right, but K_c stays the same.)

(d) The amount of COCl_2 **decreases**, because the forward reaction is exothermic, so the position moves left, in the endothermic direction, to absorb the extra heat **B1**. K_c **decreases B1**.

Example 2

Hydrogen sulfide decomposes when heated.



2.00 mol of H_2S is heated in a closed container. At equilibrium the mixture contains 0.400 mol of S_2 and the total pressure is 120 kPa.

(a) Calculate the mole fraction of each gas in the equilibrium mixture. **[2]**

(b) Calculate the partial pressure of each gas. **[1]**

(c) Write the expression for K_p and calculate its value, with units. **[3]**

(d) The total pressure is increased at constant temperature. State and explain the effect on the position of equilibrium. **[2]**

(a) 0.400 mol of S_2 forms, so 0.800 mol of H_2 forms and 0.800 mol of H_2S is used, leaving 1.20 mol. Total = 1.20 + 0.800 + 0.400 = 2.40 mol. **M1**

$$x(\text{H}_2\text{S}) = \frac{1.20}{2.40} = 0.500, x(\text{H}_2) = \frac{0.800}{2.40} = 0.333, x(\text{S}_2) = \frac{0.400}{2.40} = 0.167. \text{A1}$$

(b) $p(\text{H}_2\text{S}) = 0.500 \times 120 = 60.0 \text{ kPa}$, $p(\text{H}_2) = 40.0 \text{ kPa}$, $p(\text{S}_2) = 20.0 \text{ kPa}$. **B1** (They add up to 120 kPa, a useful check.)

$$\text{(c)} K_p = \frac{p(\text{H}_2)^2 \times p(\text{S}_2)}{p(\text{H}_2\text{S})^2} \text{B1}$$

$$K_p = \frac{40.0^2 \times 20.0}{60.0^2} = 8.89 \text{ kPa}$$

M1 for substituting partial pressures, **A1** for the value with units ($\Delta n = 3 - 2 = 1$, so kPa).

(d) The position moves to the **left B1**, because there are fewer moles of gas on the left (2) than on the right (3), so moving left lowers the pressure **B1**.

Example 3

A pink solution containing $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ ions is mixed with chloride ions in a stoppered flask. A dynamic equilibrium is set up.



$[\text{CoCl}_4]^{2-}$ is blue. The equilibrium mixture is purple.

(a) State what is meant by dynamic equilibrium. **[1]**

(b) State what is observed, and explain, when: (i) the flask is placed in iced water; **[2]** (ii) concentrated hydrochloric acid is added; **[2]** (iii) a catalyst is added. **[1]**

(a) The rates of the forward and reverse reactions are equal, and the concentrations of reactants and products stay constant. **B1**

(b)(i) The mixture becomes **more pink B1**. The forward reaction is endothermic, so lowering the temperature moves the position left, in the exothermic direction, releasing heat to oppose the cooling **B1**.

(ii) The mixture becomes **more blue B1**. The concentration of Cl^- increases, so the position moves right to use up some of the added chloride ions **B1**.

(iii) No change in colour: a catalyst speeds up the forward and reverse reactions equally, so the position does not move. **B1**

Example 4

20.0 cm³ of 0.150 mol dm⁻³ ethanoic acid is titrated with 0.100 mol dm⁻³ potassium hydroxide, added from a burette until it is in excess.

(a) Write an ionic equation for the neutralisation. **[1]**

(b) Calculate the volume of KOH needed to reach the equivalence point. **[2]**

(c) Describe the shape of the pH titration curve: the starting pH, the volume and pH range of the vertical section, and the final pH. **[3]**

(d) Use the data to choose the most suitable indicator. Explain your choice. **[2]**

Indicator	pH range of colour change
methyl orange	3.1–4.4
bromocresol green	3.8–5.4
phenolphthalein	8.3–10.0

(e) Ethanoic acid is a weak Brønsted–Lowry acid. Explain what this means and write an equation for its reaction with water. **[2]**

(a) $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$ **B1** (or $\text{CH}_3\text{COOH} + \text{OH}^- \rightarrow \text{CH}_3\text{COO}^- + \text{H}_2\text{O}$).

(b) Moles of acid = $0.0200 \times 0.150 = 0.00300$ mol **M1**. The ratio is 1 : 1, so 0.00300 mol of KOH: volume = $\frac{0.00300}{0.100} = 0.0300$ dm³ = 30.0 cm³ **A1**.

(c) Starts at about pH 3 (a weak acid) **B1**; rises slowly, then near-vertical at 30.0 cm³ from about pH 7 to 11 **B1**; levels off at about pH 12–13 (a strong alkali in excess) **B1**.

(d) Phenolphthalein **B1**. Its range (8.3–10.0) lies entirely on the vertical section (about 7–11); methyl orange and bromocresol green would change colour well before the equivalence point **B1**.

(e) It is a proton donor that is only partly dissociated in water **B1**. $\text{CH}_3\text{COOH}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{COO}^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$ **B1**.

Common mistakes

- **Defining Le Chatelier's principle as dynamic equilibrium**, or saying "the system shifts". Examiners want "the **position of equilibrium** moves to **minimise** the change".
- **Leaving out the closed system** or the constant concentrations when defining dynamic equilibrium. Give equal rates, constant concentrations and, if asked for the condition, a closed system.
- **Saying K changes because the position moved**. Pressure, concentration and catalysts move the position (or not) but never change K . Only temperature does. Examiners see this every year.
- **Saying a catalyst increases the yield**. It gives the same equilibrium, faster.
- **Adding instead of multiplying** in the K expression, or forgetting the powers. 2NH_3 means $[\text{NH}_3]^2$, and wrong powers also give wrong units.
- **Using concentrations in K_p** or square brackets around partial pressures. K_p uses p .
- **Forgetting to change the other amounts**. In a moles table, every reactant goes down and every product goes up, in the ratio of the equation: when 1 mol of N_2 reacts, 3 mol of H_2 is used and 2 mol of NH_3 forms.
- **Mixing up amounts and concentrations**. Divide moles by the volume before substituting into K_c , unless the moles on each side are equal.
- **Wrong mole fraction**. The denominator is the total moles of **gas at equilibrium**, not the starting total.
- **Counting solids or liquids** when deciding which side has fewer moles of gas.
- **Explaining a colour change without a reason**. "Goes more red" earns one mark; the second needs "because the forward reaction is exothermic" or "fewer moles of gas on the right".
- **"Releases protons"** instead of **"donates protons"** for a Brønsted–Lowry acid; and confusing a strong acid with a concentrated one.
- **Titration curve sketches** that start at the wrong pH, put the vertical section at the wrong volume, or end at the wrong level. Check all three for each curve.

How to get the marks

- **"State what is meant by..."**: a definition, word for word in its key terms: "rates of forward and reverse reactions equal", "closed system", "proton donor", "fully dissociated".
- **"State and explain"** or **"describe and explain"**: the direction or observation first, then the reason from the equation (ΔH sign, moles of gas, which substance is used up). The reason is a separate mark.
- **Observations** are colours of the **mixture**: "turns more orange", "fades", never "orange product forms".
- **"Deduce"**: work it out from the data given (a graph, a table, a change in K). Say what the data shows before your conclusion.

- **Calculations:** show every step (moles at equilibrium, then concentrations or partial pressures, then K). Each step usually earns a method mark, and an error early on is followed through.
- **Significant figures:** give K to three significant figures unless asked otherwise, and keep more figures in the middle of a calculation.
- **Units of K** are often a separate mark. Work them out from the expression, never by guessing.
- **"Sketch"** a titration curve: the right starting pH, the vertical section at the right volume and pH range, and the right final pH. Label the axes if they aren't printed.
- **Multiple choice:** check every option against the equation. Statements about K , the catalyst, units and yield are common traps.

Quick check

1. For the reaction $2\text{NO}(\text{g}) + 2\text{H}_2(\text{g}) \rightleftharpoons \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$, write the expression for K_c and give its units. What are the units of K_p in kPa?
2. For $\text{CO}(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}_2(\text{g}) + \text{H}_2(\text{g})$, $K_c = 4.00$ at a certain temperature. An equilibrium mixture contains $[\text{CO}] = 0.200 \text{ mol dm}^{-3}$ and $[\text{H}_2\text{O}] = 0.450 \text{ mol dm}^{-3}$, and $[\text{CO}_2] = [\text{H}_2]$. Calculate $[\text{CO}_2]$.
3. $\text{C}(\text{s}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + \text{H}_2(\text{g})$, $\Delta H = +131 \text{ kJ mol}^{-1}$. Write the expression for K_c with units, and state the effect on (i) the position of equilibrium of raising the pressure, (ii) K_c of raising the temperature, (iii) K_c of adding a catalyst.
4. In the reaction $\text{HNO}_3 + \text{H}_2\text{SO}_4 \rightleftharpoons \text{H}_2\text{NO}_3^+ + \text{HSO}_4^-$, identify the Brønsted–Lowry acid and base on the left, and the two conjugate pairs.
5. 25.0 cm^3 of $0.0800 \text{ mol dm}^{-3}$ sulfuric acid is titrated with $0.100 \text{ mol dm}^{-3}$ sodium hydroxide. Calculate the volume of NaOH at the equivalence point, and say whether methyl orange (3.1–4.4) or phenolphthalein (8.3–10.0) is suitable.

Answers

1. $K_c = \frac{[\text{N}_2][\text{H}_2\text{O}]^2}{[\text{NO}]^2[\text{H}_2]^2}$. $\Delta n = 3 - 4 = -1$, so the units are $(\text{mol dm}^{-3})^{-1} = \text{dm}^3 \text{ mol}^{-1}$. K_p has units of kPa^{-1} .
2. $4.00 = \frac{[\text{CO}_2]^2}{0.200 \times 0.450}$, so $[\text{CO}_2]^2 = 4.00 \times 0.0900 = 0.360$ and $[\text{CO}_2] = 0.600 \text{ mol dm}^{-3}$.
3. $K_c = \frac{[\text{CO}][\text{H}_2]}{[\text{H}_2\text{O}]}$ (carbon is a solid, so it is left out), units mol dm^{-3} . (i) Moves left: 1 mole of gas on the left, 2 on the right. (ii) K_c increases: the forward reaction is endothermic. (iii) No change: a catalyst only makes equilibrium arrive faster.
4. H_2SO_4 is the acid (it donates H^+) and HNO_3 is the base (it accepts it). Pairs: $\text{H}_2\text{SO}_4 / \text{HSO}_4^-$ and $\text{H}_2\text{NO}_3^+ / \text{HNO}_3$.
5. Moles of $\text{H}_2\text{SO}_4 = 0.0250 \times 0.0800 = 0.00200 \text{ mol}$; $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$, so 0.00400 mol of NaOH , volume = $\frac{0.00400}{0.100} = 0.0400 \text{ dm}^3 = 40.0 \text{ cm}^3$. Strong acid and strong alkali: the vertical section runs from about 3 to 11, so **both** indicators are suitable.

Past questions to try

- **9701/22/O/N/25 Q2**: part (e) defines dynamic equilibrium and predicts colour changes on warming and raising the pressure.
- **9701/24/M/J/25 Q3**: equilibrium amounts from a moles table, then an unknown amount from K_c .
- **9701/23/M/J/24 Q2**: Haber process yields, K_p with units, mole fraction and partial pressure.
- **9701/22/F/M/22 Q3**: part (d) defines a strong acid and sketches two pH titration curves.