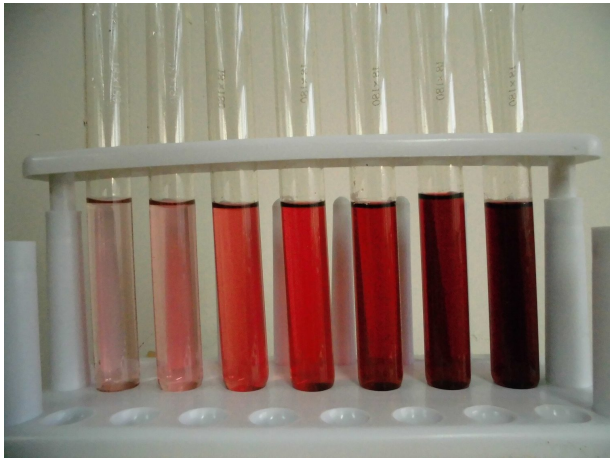
Equilibria

A-Level Chemistry

Reversible reactions and dynamic equilibrium



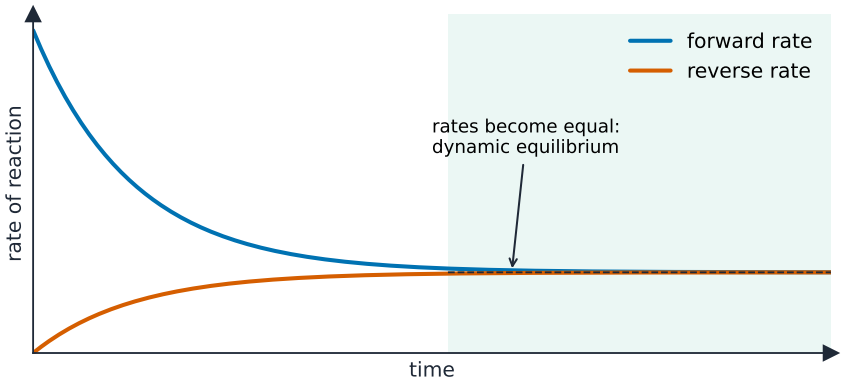
Cobalt chloride changes colour as its equilibrium shifts.

A **reversible reaction** 可逆反应 can go both ways. The **forward reaction** 正反应 makes products; the **reverse reaction** 逆反应 turns products back into reactants. We show this with the sign $\rightleftharpoons$.

In a **closed system** 封闭系统 (nothing enters or leaves), the reaction reaches a **dynamic equilibrium** 动态平衡 when:

- the rate of the forward reaction equals the rate of the reverse reaction.
- the concentrations of reactants and products stay constant.

It is called *dynamic* because both reactions are still happening — they just cancel out. A closed system is needed, or products would escape and equilibrium could never be reached.



at equilibrium: $\text{rate}_f = \text{rate}_r \cdot K_c = \frac{[\text{products}]}{[\text{reactants}]}$ (stoich. powers)

At dynamic equilibrium the forward and reverse rates have become equal, so the concentrations stay constant

Le Chatelier's principle

Le Chatelier's principle 勒夏特列原理 says: if you change a system at equilibrium, the position of **equilibrium** 平衡 moves to oppose (reduce) that change.

Change you make	Which way the equilibrium moves
increase concentration of a reactant	towards the products
increase pressure	towards the side with fewer gas molecules
increase temperature	towards the **endothermic** direction
add a **catalyst** 催化剂	no shift (it speeds up both ways equally)

A catalyst lets equilibrium be reached faster, but it does **not** change the position of equilibrium.

the equilibrium shifts to *oppose* the change you make

increase [reactant]	→ towards the **products**
increase pressure	→ to the **fewer-gas-molecules** side
increase temperature	→ in the **endothermic** direction

a **catalyst** speeds both directions equally, so it gives *no shift*

Le Chatelier's principle: the equilibrium always shifts to oppose the change you make

Equilibrium constants

For a reaction at equilibrium, the **equilibrium constant** 平衡常数 links the amounts of products and reactants. For the reaction $aA + bB \rightleftharpoons cC + dD$:

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

where the square brackets mean concentration in mol dm^{-3}.

Worked example. At equilibrium the mixture $H_2 + I_2 \rightleftharpoons 2HI$ has $[H_2] = 0.20$, $[I_2] = 0.20$ and $[HI] = 1.6$ mol dm^{-3}. Find K_c.

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{1.6^2}{0.20 \times 0.20} = 64$$

(no units here, because the concentration powers cancel top and bottom).

Worked example (ICE table). 1.00 mol each of H_2 and I_2 are sealed in a 1.00 dm^3 flask; at equilibrium 0.20 mol of H_2 remains. Set out an **ICE table** (Initial, Change, Equilibrium):

	H_2	I_2	$2HI$
Initial	1.00	1.00	0
Change	−0.80	−0.80	+1.60
Equilibrium	0.20	0.20	1.60

0.80 mol of H_2 reacted, so $2 \times 0.80 = 1.60$ mol HI formed. In the 1.00 dm^3 flask the concentrations equal the moles, so $K_c = \dfrac{1.60^2}{0.20 \times 0.20} = 64$.

For reactions of gases, we use **partial pressure** 分压 instead of concentration. The partial pressure of a gas is the share of the total pressure that it provides. It is found from the **mole fraction** 摩尔分数 (the fraction of all the moles that are that gas):

$$\text{partial pressure} = \text{mole fraction} \times \text{total pressure}$$

The constant written with partial pressures is K_p.

Worked example. A gas mixture holds 2.0 mol of N_2 and 6.0 mol of H_2 at a total pressure of 200 kPa. Find the partial pressure of each gas.

There are 8.0 mol in total, so

$$p(N_2) = \frac{2.0}{8.0} \times 200 = 50 \text{ kPa}, \qquad p(H_2) = \frac{6.0}{8.0} \times 200 = 150 \text{ kPa}.$$

Worked example (a numeric K_p). For $N_2 + 3H_2 \rightleftharpoons 2NH_3$ the equilibrium partial pressures are $p(N_2) = 20$, $p(H_2) = 40$ and $p(NH_3) = 10$ kPa. Then

$$K_p = \frac{p(NH_3)^2}{p(N_2)\,p(H_2)^3} = \frac{10^2}{20 \times 40^3} = 7.8 \times 10^{-5} \text{ kPa}^{-2},$$

the units coming from $\dfrac{\text{kPa}^2}{\text{kPa} \times \text{kPa}^3} = \text{kPa}^{-2}$.

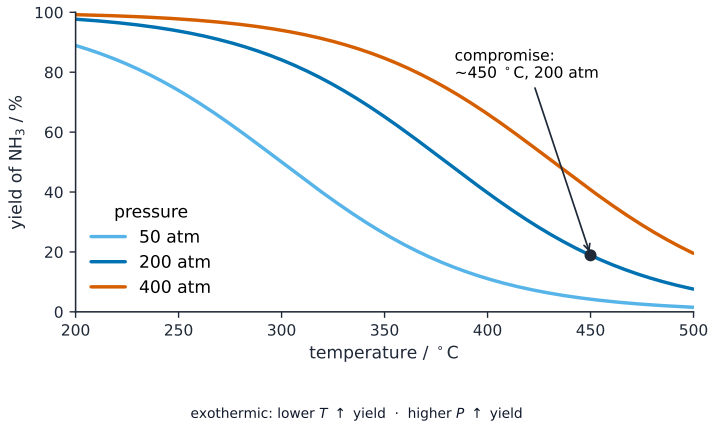
Only **temperature** changes the value of K_c or K_p. Changing concentration or pressure shifts the position of equilibrium but leaves the constant unchanged; adding a catalyst changes **neither** the constant nor the position — it only makes equilibrium arrive faster.

The Haber and Contact processes

Ammonia for fertiliser is made by the Haber process in large industrial plants like this one.

These two industrial processes are chosen by balancing yield, rate and cost using Le Chatelier's principle.

- the **Haber process** 哈伯法 makes ammonia: $N_2 + 3H_2 \rightleftharpoons 2NH_3$ (exothermic). It uses about 450 °C, 200 atm and an iron catalyst. A low temperature would give more ammonia but too slowly, so a moderate temperature is a compromise.
- the **Contact process** 接触法 makes sulfur trioxide: $2SO_2 + O_2 \rightleftharpoons 2SO_3$ (exothermic). It uses about 450 °C, near 1–2 atm and a vanadium(V) oxide catalyst.

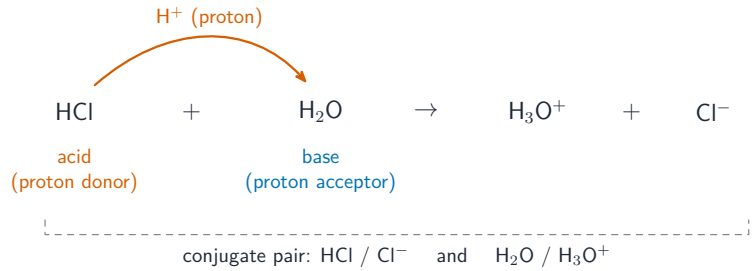


The Haber equilibrium gives more ammonia at lower temperature and higher pressure; about 450 °C and 200 atm is the working compromise

Acids and bases: the Brønsted–Lowry theory

A **proton** 质子 is simply an H^+ ion. The Brønsted–Lowry theory defines acids and bases by what they do with protons:

- an **acid** 酸 is a **proton donor** 质子供体 (it gives away H^+).
- a **base** 碱 is a **proton acceptor** 质子受体 (it takes H^+). A base that dissolves in water is called an alkali.



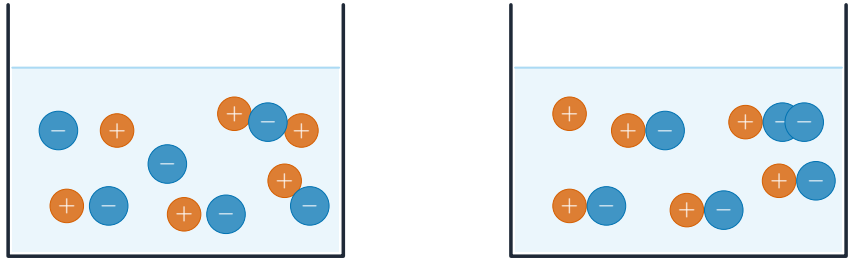
Brønsted–Lowry: the acid donates a proton (H^+) to the base, making two conjugate acid–base pairs

Common acids you must know: hydrochloric acid (HCl), sulfuric acid (H_2SO_4), nitric acid (HNO_3) and ethanoic acid (CH_3COOH). Common alkalis: sodium hydroxide (NaOH), potassium hydroxide (KOH) and ammonia (NH_3).

Strong and weak acids and bases

This is about how fully an acid or base splits up in water — not how concentrated it is.

- a **strong acid** 强酸 or **strong base** 强碱 is fully **dissociated** 解离 in water (almost every molecule splits into ions).
- a **weak acid** 弱酸 or **weak base** 弱碱 is only partly dissociated (most molecules stay whole).



strong acid
fully dissociated: many ions

weak acid
mostly molecules: few ions

Same concentration, different ionisation: a strong acid is fully dissociated into ions; a weak acid stays mostly as whole molecules

The **pH** scale measures how acidic a solution is: pure water is pH 7, acids are below 7, and alkalis are above 7. A strong acid has a lower pH than a weak acid of the same concentration.

You can tell strong and weak acids apart by:

- **reaction with a reactive metal**: a strong acid fizzes faster.
- **pH**: measured with a pH meter or **universal indicator** 通用指示剂.
- **electrical conductivity**: a strong acid conducts better, because it has more ions.

Neutralisation, salts and titration curves

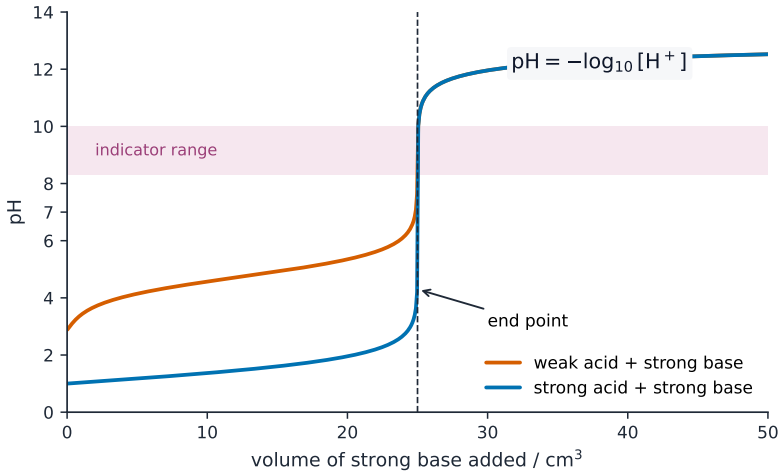
Neutralisation 中和 happens when the hydrogen ions from an acid react with the hydroxide ions from an alkali:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

A **salt** 盐 is also formed, from the rest of the acid and base.

In a **titration** 滴定 you add one solution to another and follow the pH. The **titration curve** 滴定曲线 has a steep, almost vertical jump near the end point. The exact shape depends on whether each reactant is strong or weak.

To pick an **indicator** 指示剂, choose one whose colour change falls inside that steep jump. For a strong acid with a strong base most indicators work; for a weak acid with a strong base you need one that changes in the higher pH range.



Titration curves each have a steep pH jump at the end point. The weak-acid jump sits higher, so the indicator must change colour in that range

Exam tips

- Define **dynamic equilibrium** with all three points: forward and reverse rates equal, concentrations constant, closed system.
- Answer **Le Chatelier** questions by stating the change, the direction of shift, and the reason; a **catalyst does not shift** the position (it speeds both rates).
- Write K_c as **products over reactants**, each raised to its balancing number; include state symbols to decide what appears.
- Justify Haber/Contact conditions as a **compromise** between yield, rate and cost — do not just quote them.
- Brønsted: acid = **proton donor**, base = proton acceptor; conjugate pairs differ by one H^+.