

ESAT

Chemistry

Equations and equilibrium

Free lesson notes

Covers

C3.1 to C3.5

Format

Lesson, worked examples, self-check

Price

Free

About this lesson

A chemical equation is a very compact piece of chemistry. It tells you what reacts, what is made, in what proportions and in what state, and it obeys one iron rule: atoms are never created or destroyed, only rearranged. Nearly every quantitative question in chemistry begins by writing a correct balanced equation, so this topic is the foundation for moles, titrations, yields and electrolysis.

We begin with the idea that a reaction rearranges atoms, and then learn how to write the *formulae* that go into an equation, first for ionic compounds (by balancing charges) and then for covalent ones. We add state symbols, and learn to balance equations systematically. Then we go one level deeper: ionic equations, which show only the particles that actually change, and half-equations, which show the electrons being transferred. The last two sections cover reversible reactions: what a dynamic equilibrium is, and how changing concentration, pressure and temperature shifts it.

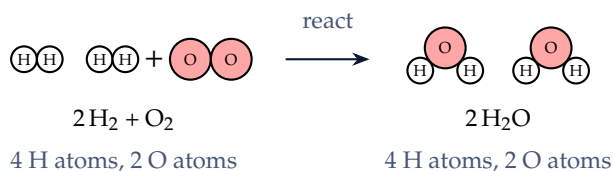
How to use it. Cover each worked example and try it before reading the solution. There is no calculator in the real ESAT, so everything here is done by hand. Finish with the self-check, then the practice set.

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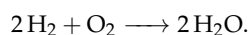
1 Reactions rearrange atoms (C3.1)

In a chemical reaction the substances you start with, the **reactants**, are changed into new substances, the **products**. The atoms themselves are exactly the same atoms; they simply end up joined to different partners. Bonds in the reactants are broken and new bonds form, so the atoms are rearranged. Because no atoms are gained or lost, the total mass of the products equals the total mass of the reactants: this is the **conservation of mass**. (If a reaction in an open container seems to lose mass, it is because a gas has escaped; if it seems to gain mass, a gas, such as oxygen from the air, has been added.)

A **word equation** names the substances: *hydrogen + oxygen → water*. A **symbol equation** uses formulae, and has to be **balanced**: the same number of each kind of atom on both sides.



In the picture, two hydrogen molecules (each of two H atoms) react with one oxygen molecule (two O atoms). There are 4 hydrogen atoms and 2 oxygen atoms before, and 4 and 2 afterwards, shared out as two molecules of water, each with two H and one O. The balanced equation is



The big numbers in front of formulae are **coefficients**: they say how many of each particle take part. The small numbers within a formula (subscripts) are part of the identity of the substance: H_2O is water, but H_2O_2 is a different substance (hydrogen peroxide).

Worked example: conservation of mass

24 g of magnesium burns in air and forms 40 g of magnesium oxide. What mass of oxygen has combined with the magnesium?

Mass is conserved, so the oxygen supplied is $40 - 24 = 16$ g. (The product is heavier than the magnesium because oxygen atoms have been added.)

Common mistakes

- Changing a subscript to balance an equation. Subscripts are fixed by the substance; only coefficients may be changed.
- Thinking atoms are used up. They are all still there, rearranged.

- Forgetting the diatomic elements: hydrogen, nitrogen, oxygen and the halogens exist as H_2 , N_2 , O_2 , F_2 , Cl_2 , Br_2 , I_2 .

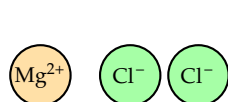
2 Formulae of ionic compounds (C3.2)

An ionic compound is made of positive and negative **ions**. Overall it has no charge, so the positive and negative charges must cancel. This one rule tells you its formula. First you need the charges of the common ions. Metals lose electrons to become positive ions, non-metals gain electrons to become negative ions, and the charge follows from the group:

Group 1	Group 2	Aluminium	Group 16	Group 17
Li^+ , Na^+ , K^+	Mg^{2+} , Ca^{2+}	Al^{3+}	O^{2-} , S^{2-}	F^- , Cl^- , Br^- , I^-

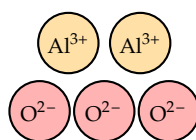
compound positive ions	compound negative ions
H^+ (hydrogen), NH_4^+ (ammonium)	OH^- , NO_3^- , CO_3^{2-} , SO_4^{2-}

Method. Write the ions with their charges, then choose the smallest numbers of each that make the total charge zero.



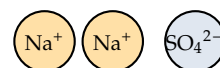
$$(+2) + 2 \times (-1) = 0$$

$MgCl_2$



$$2 \times (+3) + 3 \times (-2) = 0$$

Al_2O_3



$$2 \times (+1) + (-2) = 0$$

Na_2SO_4

- Magnesium chloride: Mg^{2+} needs two Cl^- to balance, so $MgCl_2$.
- Aluminium oxide: Al^{3+} and O^{2-} . The smallest common multiple of 3 and 2 is 6: two Al^{3+} (+6) and three O^{2-} (-6), so Al_2O_3 .
- When a compound ion appears more than once, use brackets: calcium nitrate is Ca^{2+} with two NO_3^- , so $Ca(NO_3)_2$ (not $CaNO_{32}$). Ammonium sulfate is $(NH_4)_2SO_4$.

Roman numerals. Transition metals can form more than one ion, so the charge is given in brackets in the name: iron(II) chloride contains Fe^{2+} and is $FeCl_2$; iron(III) chloride contains Fe^{3+} and is $FeCl_3$; copper(II) oxide is CuO .

Covalent compounds. Atoms of non-metals share electrons. Here the formula follows from how many bonds each atom makes: hydrogen and the halogens make 1, oxygen 2, nitrogen 3, carbon 4. So H_2O , NH_3 , CH_4 , HCl , and CO_2 (carbon forms two double bonds). Names help too: "di", "tri" and "tetra" tell you the number of atoms, as in carbon *dioxide* or phosphorus *trichloride*. A name ending in "-ide" means only two elements are present; "-ate" means oxygen is present as well (sulfate, nitrate, carbonate).

Worked example: formulae from charges

Write the formulae of (a) sodium oxide, (b) iron(III) oxide, (c) magnesium nitrate.

(a) Na^+ and O^{2-} : two sodium ions per oxide, Na_2O . (b) Fe^{3+} and O^{2-} : $2 \times 3 = 3 \times 2$, so Fe_2O_3 . (c) Mg^{2+} and NO_3^- : two nitrates, $Mg(NO_3)_2$.

Common mistakes

- Writing the charges as subscripts (Mg_2Cl). Charges decide the *numbers* of ions; the formula has no charges.
- Forgetting brackets around a compound ion that appears more than once.
- Cancelling to the wrong ratio. The formula of an ionic compound is the simplest whole-number ratio of ions, so sodium oxide is Na_2O , not NaO .

3 State symbols (C3.3)

Equations often show the physical state of each substance, in brackets after the formula:

symbol	meaning
(s)	solid
(l)	liquid
(g)	gas
(aq)	aqueous: dissolved in water

For example $\text{Mg(s)} + 2\text{HCl(aq)} \longrightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)}$ says that solid magnesium reacts with a solution of hydrochloric acid to give a solution of magnesium chloride and bubbles of hydrogen gas. Do not confuse (l) with (aq): pure water is $\text{H}_2\text{O(l)}$, whereas salt dissolved in water is NaCl(aq) . The symbols matter when you write ionic equations (next sections) because only dissolved ionic compounds split into ions.

Worked example: reading state symbols

In $\text{CaCO}_3\text{(s)} + 2\text{HCl(aq)} \longrightarrow \text{CaCl}_2\text{(aq)} + \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$, what would you see?

A solid (calcium carbonate, e.g. marble chips) disappears into a colourless solution as bubbles of carbon dioxide gas are given off. The water produced is a liquid.

Common mistakes

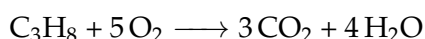
- Writing (aq) for a pure liquid such as water. Water formed in a reaction is (l).
- Writing (g) or (l) for an insoluble solid that forms in a solution (a precipitate is (s)).

4 Balancing equations (C3.4)

To balance an equation: (1) write the correct formula for every substance, and *never change them afterwards*; (2) count the atoms of each element on each side; (3) change coefficients, one element at a time, until every count matches; (4) check the final tally.

A good order. Balance elements that appear in only one substance on each side first; leave hydrogen and oxygen (which often appear in several substances) until last; and if oxygen then comes out as an odd number against an even one, double everything.

Worked through: the combustion of propane. Unbalanced: $\text{C}_3\text{H}_8 + \text{O}_2 \longrightarrow \text{CO}_2 + \text{H}_2\text{O}$. Carbon: 3 on the left, so 3 CO_2 . Hydrogen: 8 on the left, so 4 H_2O . Oxygen on the right: $3 \times 2 + 4 \times 1 = 10$, so 5 O_2 on the left.



	C	H	O
left	3	8	10
right	3	8	10

same number of each atom
on both sides: balanced

Odd and even. For ethane, $\text{C}_2\text{H}_6 + \text{O}_2 \longrightarrow \text{CO}_2 + \text{H}_2\text{O}$: carbon gives 2 CO_2 and hydrogen gives 3 H_2O . Oxygen on the right is $4 + 3 = 7$, which is odd, but oxygen on the left comes in pairs. The remedy is to double everything to start with: $2\text{C}_2\text{H}_6 + 7\text{O}_2 \longrightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$. Do not use a half ($3\frac{1}{2}\text{O}_2$) in an ESAT answer unless it is asked for.

Worked example: balancing

Balance $\text{Fe}_2\text{O}_3 + \text{CO} \longrightarrow \text{Fe} + \text{CO}_2$.

Iron: 2 Fe on the right. Oxygen: left has $3 + 1 = 4$ if one CO; try 3 CO and 3 CO_2 : oxygen left $3 + 3 = 6$, right $3 \times 2 = 6$; carbon $3 = 3$. So $\text{Fe}_2\text{O}_3 + 3\text{CO} \longrightarrow 2\text{Fe} + 3\text{CO}_2$.

Worked example: balancing with a bracket

Balance $\text{Al} + \text{H}_2\text{SO}_4 \longrightarrow \text{Al}_2(\text{SO}_4)_3 + \text{H}_2$.

Aluminium: 2 Al. Sulfate: three SO_4 , so 3 H_2SO_4 . Hydrogen: $3 \times 2 = 6$ on the left gives 3 H_2 on the right.
 $2 \text{Al} + 3 \text{H}_2\text{SO}_4 \longrightarrow \text{Al}_2(\text{SO}_4)_3 + 3 \text{H}_2$.

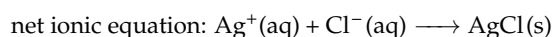
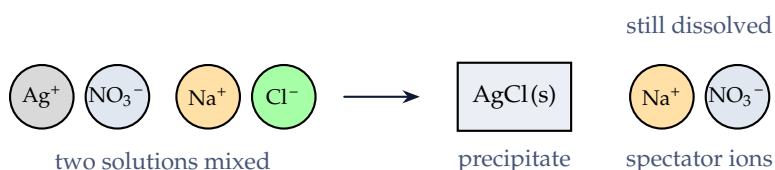
Common mistakes

- Changing a subscript to balance. CO_2 must stay CO_2 .
- Balancing oxygen first and then upsetting it when you balance hydrogen. Leave O and H to last, and recount everything at the end.
- Counting atoms in a bracket once: $\text{Ca}(\text{OH})_2$ has 2 O and 2 H; $\text{Al}_2(\text{SO}_4)_3$ has 12 O.

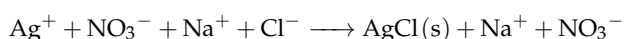
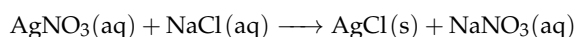
5 Ionic equations and half-equations (C3.4)

Ionic equations. When ionic compounds are dissolved in water, they exist as separate ions. In many reactions, some of those ions take no part at all; they are **spectator ions**. An ionic equation leaves them out and shows only the particles that change. To write one:

1. Write the balanced equation with state symbols.
2. Split every (aq) ionic compound into its ions. Leave solids, liquids, gases and molecular substances (such as water) as they are.
3. Cancel the ions that appear unchanged on both sides (the spectators).
4. Check that the atoms *and* the *total charge* balance.



Silver nitrate and sodium chloride solutions give a white precipitate of silver chloride:



Cancel Na^+ and NO_3^- : $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \longrightarrow \text{AgCl}(\text{s})$.

Neutralisation has the same ionic equation whatever the acid and alkali (strong ones): $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \longrightarrow \text{H}_2\text{O}(\text{l})$.

Metal displacement is also ionic: $\text{Zn}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \longrightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu}(\text{s})$.

Charge must balance too. In the displacement above, the total charge on the left is $0 + 2 = +2$ and on the right is $+2 + 0 = +2$.

Half-equations. In a reaction in which electrons are transferred (see the redox lesson), we can split the ionic equation into two halves, one showing a substance *losing* electrons and one *gaining* them. Electrons are written as e^- . For the zinc and copper reaction:



Each half-equation must balance in atoms and in charge: the left side of the second has charge $+2 + 2(-1) = 0$, the right side 0. The electrons lost in one half equal the electrons gained in the other, so adding the halves cancels them to give the ionic equation.

Worked example: net ionic equation

Barium chloride solution and sodium sulfate solution react to give a white precipitate of barium sulfate. Write the ionic equation and name the spectator ions.

Full: $\text{BaCl}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \longrightarrow \text{BaSO}_4(\text{s}) + 2\text{NaCl}(\text{aq})$. Ions: $\text{Ba}^{2+} + 2\text{Cl}^- + 2\text{Na}^+ + \text{SO}_4^{2-} \longrightarrow \text{BaSO}_4(\text{s}) + 2\text{Na}^+ + 2\text{Cl}^-$. Cancelling gives $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \longrightarrow \text{BaSO}_4(\text{s})$. The spectators are Na^+ and Cl^- .

Worked example: balancing charge

Find x in the balanced ionic equation $\text{Fe} + x \text{Ag}^+ \longrightarrow \text{Fe}^{2+} + x \text{Ag}$ given that all the atoms already balance. Total charge left = $0 + x(+1) = x$. Right = $+2 + 0 = +2$. So $x = 2$: $\text{Fe} + 2 \text{Ag}^+ \longrightarrow \text{Fe}^{2+} + 2 \text{Ag}$.

Common mistakes

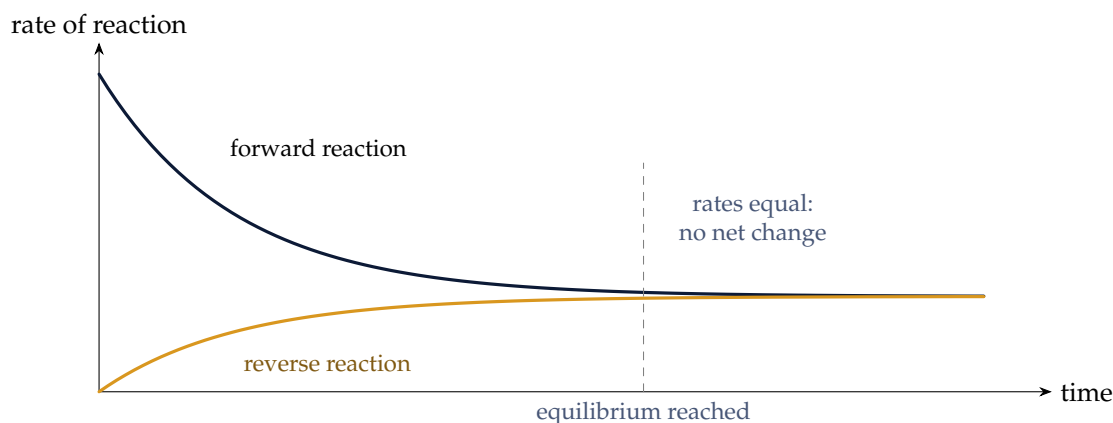
- Splitting solids or gases (or water) into ions. Only dissolved ionic compounds split.
- Forgetting to balance charge: a balanced ionic equation has the same total charge on both sides.
- Cancelling coefficients wrongly: if 2Cl^- appear on both sides, cancel both.

6 Reversible reactions and dynamic equilibrium (C3.5)

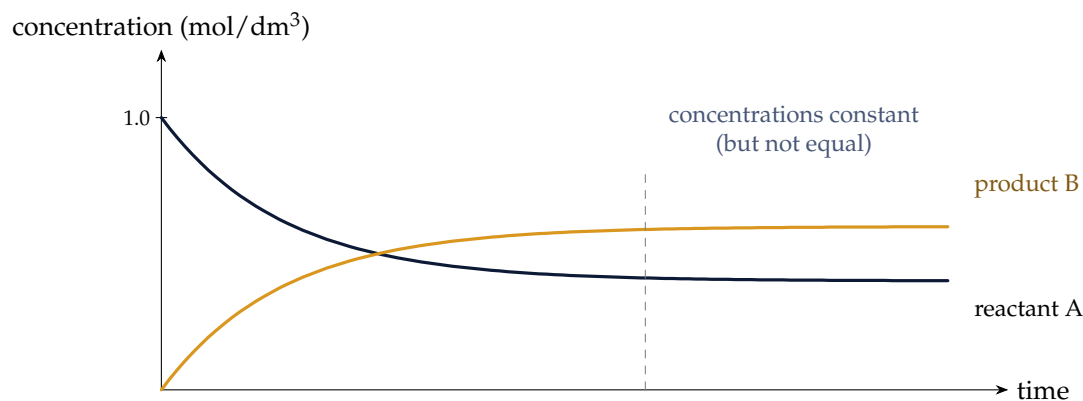
Many reactions go only one way, but some can go in both directions. These are **reversible reactions**, shown with a double arrow $\rightleftharpoons$. Examples: heating blue hydrated copper(II) sulfate drives off water to give white anhydrous copper(II) sulfate, and adding water turns it blue again, and the brown gas nitrogen dioxide forms a colourless gas dinitrogen tetroxide:



Put a reversible reaction in a *closed* container (nothing can escape) and something remarkable happens. At the start only reactants are present, so the forward reaction runs quickly. As products build up, the reverse reaction begins, and it speeds up as the products become more concentrated, while the forward reaction slows as the reactants are used up.



Eventually the two rates become *equal*. From then on, reactants are turned into products as fast as products are turned back into reactants, and the concentrations stop changing. This is **dynamic equilibrium**: “dynamic” because both reactions are still going on, and “equilibrium” because there is no overall change. It needs a closed system.



Two things to avoid. At equilibrium the concentrations are *constant*, but not necessarily *equal*: here $\text{A} \rightleftharpoons \text{B}$ settles with 0.4 and 0.6 mol/dm³ (the curves crossing is not the equilibrium). And the reaction has not stopped: the particles keep reacting.

The **position of equilibrium** describes how far the reaction has gone: if the mixture contains mostly products, it lies to the right; if mostly reactants, it lies to the left. A catalyst makes both rates faster equally, so equilibrium is reached sooner, but the position does *not* change.

Worked example: reading an equilibrium graph

In the graph above, A starts at 1.0 mol/dm^3 and B at 0, and at equilibrium A has fallen to 0.4 mol/dm^3 . What percentage of A has been converted to B?

$A \rightarrow B$ is one-to-one, so $B = 1.0 - 0.4 = 0.6$. The percentage converted is $0.6/1.0 = 60\%$.

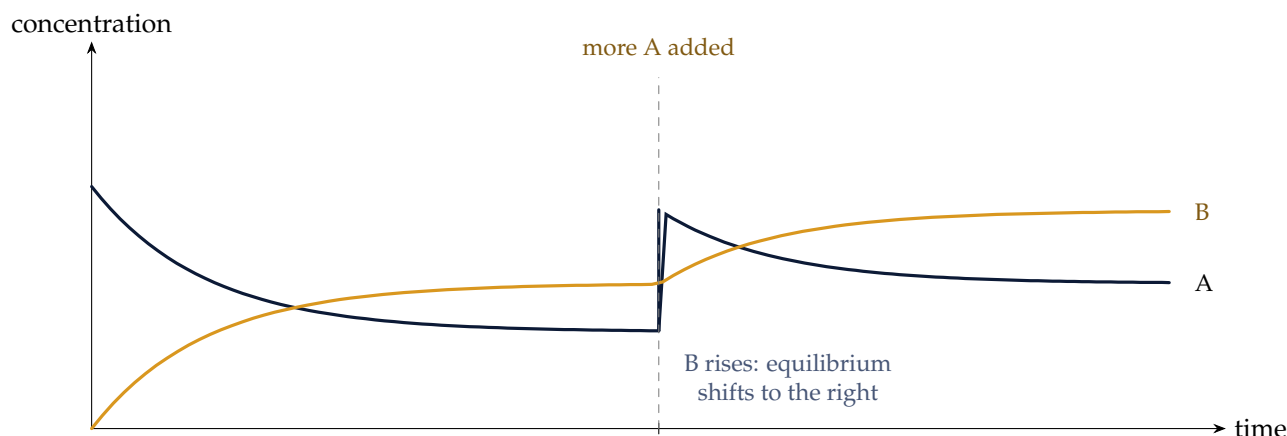
Common mistakes

- Saying the reaction stops at equilibrium. It does not: both directions continue at the same rate.
- Saying concentrations are equal at equilibrium. They are constant, not equal.
- Thinking equilibrium can be reached in an open container. Gases would escape: the system must be closed.

7 Shifting the equilibrium (C3.5)

Chemists want to push equilibria towards the products they want. The guide is **Le Chatelier's principle**: *if a change is made to a system at equilibrium, the system shifts to oppose (partly undo) the change*. We apply it to three changes.

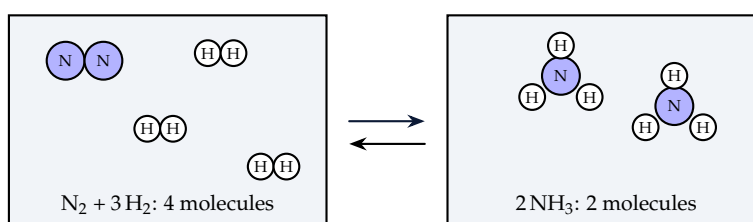
Concentration. Add more reactant: the system uses it up by making more product, so the position shifts to the *right*. Remove product (as it forms): the system makes more to replace it, again shifting right. The reverse changes shift it left.



In the graph more of A is added at 10 s. Its concentration jumps up, then falls as it is converted, while the concentration of B rises to a new, higher plateau.

Pressure (only matters if gases are involved). Raising the pressure favours the side with *fewer molecules of gas*, because that reduces the pressure. In $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$ there are 4 molecules on the left and 2 on the right, so high pressure shifts the position right (more ammonia). If the numbers of gas molecules are equal on both sides, pressure has no effect on the position.

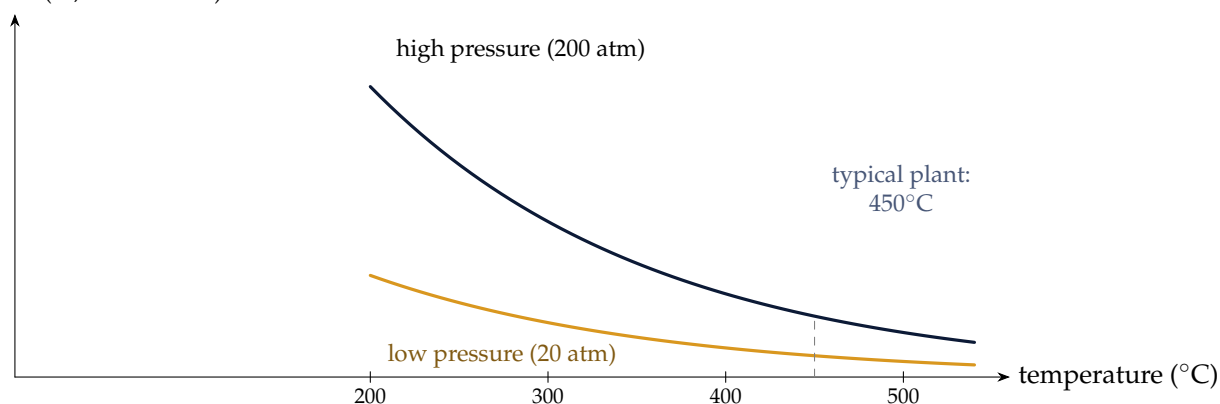
higher pressure favours the side with fewer gas molecules



Temperature. A reversible reaction is exothermic in one direction (releases heat) and endothermic in the other (takes in heat) by the same amount. Raising the temperature favours the *endothermic* direction (it absorbs the extra heat); lowering

it favours the *exothermic* direction. The Haber process reaction is exothermic to the right, so a higher temperature gives less ammonia at equilibrium.

NH₃ at equilibrium (% , illustrative)



Why the real conditions are a compromise. Low temperature gives the best yield, but the reaction is far too slow, and high pressure is expensive to maintain. The industrial Haber process therefore uses about 450 °C, about 200 atmospheres, and an iron catalyst, which speeds up the approach to equilibrium without changing the position.

change	shifts towards	note
add reactant / remove product	products (right)	
increase pressure	fewer gas molecules	no effect if equal numbers
increase temperature	endothermic direction	
add a catalyst	<i>no shift</i>	reaches equilibrium faster

Worked example: predicting the shift

For $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$ (exothermic to the right), what would you do to increase the equilibrium yield of SO₃?

Gas molecules: 2 + 1 = 3 on the left, 2 on the right, so *increase the pressure*. The forward reaction is exothermic, so *decrease the temperature*. (Removing SO₃ as it forms, or adding more SO₂ or O₂, would also help.) A catalyst would not change the yield.

Worked example: colour change in a syringe

A sealed syringe holds the equilibrium mixture $2\text{NO}_2 \rightleftharpoons \text{N}_2\text{O}_4$ (brown $\rightleftharpoons$ colourless). The plunger is pushed in quickly. What is seen?

Pushing in the plunger increases the pressure, which also squeezes the same gas into a smaller volume, so the brown colour first looks *darker*. Then the equilibrium shifts to the side with fewer molecules, the colourless N₂O₄, so the colour fades again, but remains darker than at the start.

Exam tip

For any shift question ask three things in order: (1) is it concentration, pressure or temperature? (2) for pressure, count gas molecules each side; for temperature, which direction is exothermic? (3) a catalyst never changes the position, only the rate.

Common mistakes

- Saying a catalyst increases the yield. It only speeds up how fast equilibrium is reached.
- Applying the pressure rule to solids and liquids. Count only gas molecules.
- Mixing up temperature: raising temperature favours the *endothermic* direction, always.

Check yourself

Try these without looking back, then compare with the answers on the next page.

1. Write the formula of calcium chloride and of aluminium sulfate.
2. Balance $\text{Mg} + \text{O}_2 \longrightarrow \text{MgO}$.
3. What does (aq) mean?
4. Balance $\text{C}_2\text{H}_4 + \text{O}_2 \longrightarrow \text{CO}_2 + \text{H}_2\text{O}$.
5. Write the ionic equation for neutralisation of a strong acid by a strong alkali.
6. Complete the half-equation $\text{Mg} \longrightarrow \text{Mg}^{2+} + ?$ (what must be added to balance the charge?).
7. What two things are true of a system at dynamic equilibrium?
8. For $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$ (exothermic to the right), does a higher temperature increase or decrease the yield of ammonia?

Answers to the self-check

1. CaCl_2 (one Ca^{2+} , two Cl^-); $\text{Al}_2(\text{SO}_4)_3$ ($2 \times 3 = 3 \times 2$).
2. $2\text{Mg} + \text{O}_2 \longrightarrow 2\text{MgO}$.
3. Dissolved in water (an aqueous solution).
4. $\text{C}_2\text{H}_4 + 3\text{O}_2 \longrightarrow 2\text{CO}_2 + 2\text{H}_2\text{O}$.
5. $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \longrightarrow \text{H}_2\text{O}(\text{l})$.
6. $\text{Mg} \longrightarrow \text{Mg}^{2+} + 2\text{e}^-$.
7. Forward and reverse rates are equal, and the concentrations are constant; the system is closed.
8. Decrease: higher temperature favours the endothermic (left-hand) direction.

Next step

Heurisco's *ESAT Chemistry: Equations and Equilibrium Practice Set* has twenty original exam-style questions on this lesson, with particle diagrams and concentration graphs, in order of difficulty, with a full worked solution and a note on the trap in each.