

# Chemical Reactions

IGCSE Chemistry ■ Topic 6

IGCSE Chemistry



## What we will cover

- 1 Physical vs chemical change
- 2 Collision theory & rate
- 3 Catalysts & measuring rate
- 4 Reversible reactions
- 5 Equilibrium & industry
- 6 Redox



# Physical vs chemical change



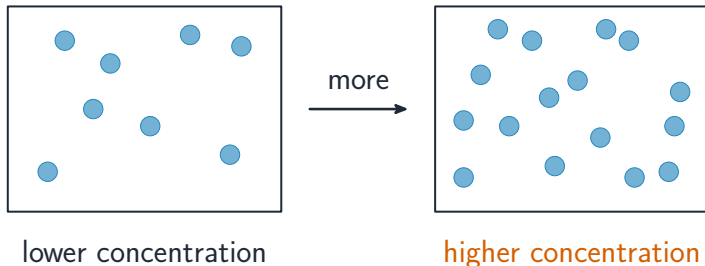
*Iron rusting – a new substance forms*

- A **physical change** makes **no new substance** – only the state or shape changes, and it usually reverses. Melting ice, dissolving sugar.
- A **chemical change** makes **one or more new substances** and is usually hard to reverse. Burning, rusting.
- Signs of a chemical change: a colour change, a gas given off, an energy change, or a **precipitate** forming.

## Collision theory and rate

**Collision theory** : particles must collide with enough energy ( $\geq E_a$ ). Faster when collisions are **more frequent** or **more energetic**:

- more **concentration** / pressure
- more **surface area**
- higher temperature



# What a collision needs

## Rate of reaction

The **rate of reaction** is how fast the **reactants** are used up, or the **products** formed.

## Frequency

More collisions per second – higher concentration, higher pressure, larger surface area – means a faster rate.

## Energy

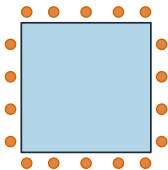
A collision only works if the particles have enough **kinetic energy** to react. Heating raises the average, so far more collisions succeed.

## The other half

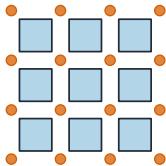
Temperature therefore acts twice, which is why it changes rate more sharply than concentration does.

Every rate factor is either *more collisions* or *more successful ones*. Say which, and the explanation is complete.

# Surface area



one big lump  
small surface  
area  $\rightarrow$  slower

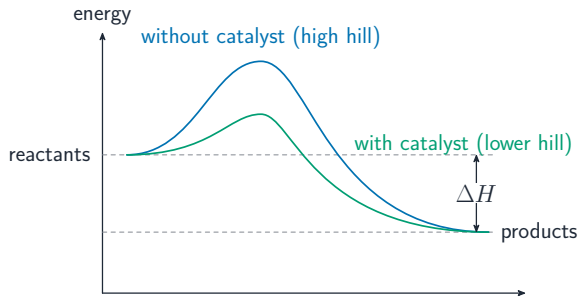


same mass  
as powder  
large surface  
area  $\rightarrow$  faster

A powder has far more surface than one lump, so acid touches **many more** particles.

More exposed particles  $\Rightarrow$   
more collisions  $\Rightarrow$  a faster  
rate.

# Catalysts

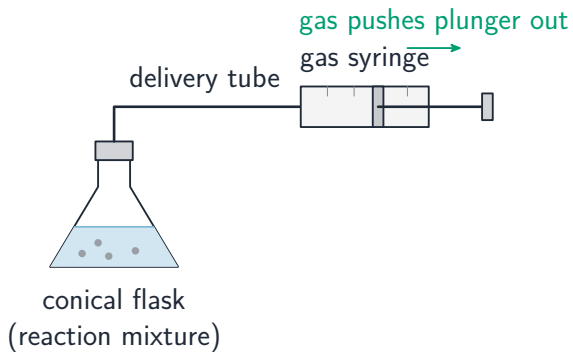


a catalyst lowers the activation energy;  $\Delta H$  is unchanged

A **catalyst** speeds a reaction but is **not used up**.

- it lowers the **activation energy**
- $\Delta H$  is unchanged
- **enzymes** are biological catalysts

## Measuring the rate



Follow a reaction by:

- loss of **mass** on a balance
- **gas volume** in a **syringe**
- time for a **precipitate** to hide a mark

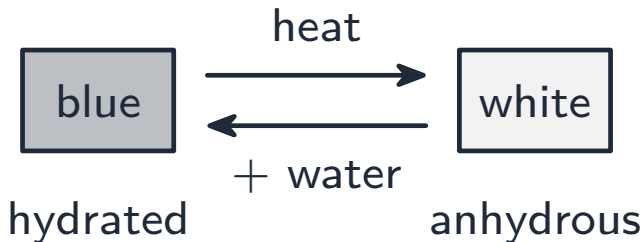
The graph is steepest at the start, then flat.



## Reversible reactions

A **reversible reaction** ( $\rightleftharpoons$ ) can go both ways.

Blue **hydrated**  $\text{CuSO}_4 \rightleftharpoons$  white **anhydrous**  $\text{CuSO}_4 + \text{water}$ . Adding water turns it blue – a test for water.



# A reaction that runs both ways

## Reverse reaction

In a reversible reaction the products re-form the reactants – the **reverse reaction** – and both run at once.

## Equilibrium

At equilibrium the two rates are *equal*, so the amounts stop changing. Nothing has stopped happening.

## A classic test

Hydrated copper(II) sulfate turns white when heated and blue again with water – reversible, and visible.

## Redox in one line

A **redox reaction** is any reaction where electrons transfer: one species is oxidised as another is reduced.

Equilibrium is dynamic. Marks are lost by writing that the reaction “stops”.

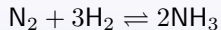
# Equilibrium

At **equilibrium** (closed system): rate forward = rate reverse  
concentrations stop changing (but the reaction is still going)

To shift the position: **heat** → endothermic side; **pressure** → fewer-molecules side; **add a substance** → the other side. A **catalyst does not** shift it.

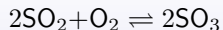
## Industry: Haber and Contact

### Haber – ammonia



450°C, ~200 atm, iron catalyst

### Contact – SO<sub>3</sub>



450°C, ~2 atm, V<sub>2</sub>O<sub>5</sub> catalyst

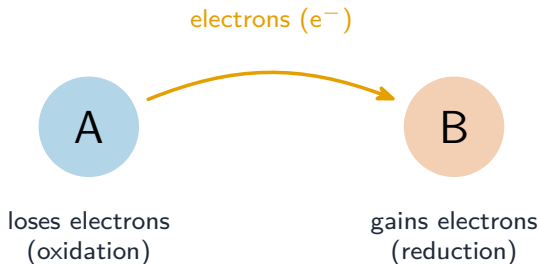
Conditions are a **compromise** : medium pressure (safe), fairly high temperature (good rate).

## Why these conditions are chosen

- A higher pressure would give more product, but very high pressure is dangerous and expensive, so a medium pressure is used.
- A lower temperature would give more product (both forward reactions are exothermic), but the reaction would be too slow, so a fairly high temperature is used to keep a good rate.
- The catalyst speeds up the reaction without changing the position of equilibrium, which lowers cost.



# Redox



**Oxidation & reduction** happen together.

**OIL RIG:** Oxidation Is Loss, Reduction Is Gain – of electrons.

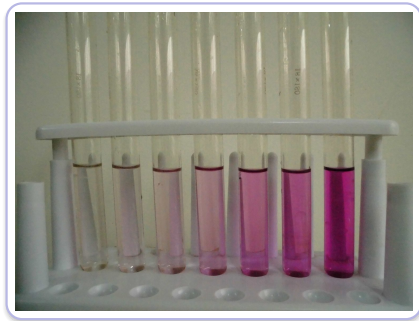
An **oxidising agent** is itself reduced (purple  $\text{MnO}_4^-$  fades).

## Oxidation number rules

The oxidation number is shown by a Roman numeral, as in iron(II) and iron(III).

- An element that is not combined has an oxidation number of 0.
- A single-atom ion has an oxidation number equal to its charge (so  $\text{Na}^+$  is +1).
- The oxidation numbers in a compound add up to 0.
- The oxidation numbers in an ion add up to the charge on the ion.

## Oxidising and reducing agents



Potassium manganate(VII) is deep purple; as it is reduced (or diluted) the purple fades to colourless

## A test you can see

### The reaction

Chlorine displaces iodine from **potassium iodide** :  $\text{Cl}_2 + 2\text{KI} \rightarrow 2\text{KCl} + \text{I}_2$ .

### What you see

The solution turns brown as **iodine** is released – the more reactive halogen has taken the electrons.

### Naming it

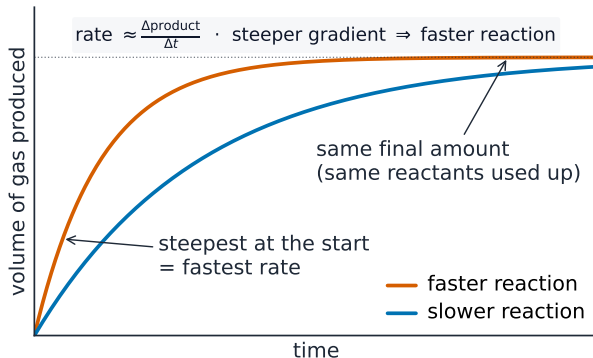
Chlorine is the oxidising agent (it gains electrons and is reduced); iodide is the reducing agent.

### Also on the list

Reactions with **oxygen** are the everyday oxidation: burning, rusting, and the tarnishing of a fresh metal surface.

An oxidising agent is the species that is *reduced*. The two words describe the same electron moving.

## Measuring the rate, continued



The rate is fastest at the start (steepest) and the line levels off when the reaction finishes

# Measuring a rate, and the two industrial processes

**change in mass**

stand the flask on a balance – if a gas escapes, the **mass falls**.  
Record at regular times

**volume of gas**

collect it in a gas **syringe** and read the volume at regular times

**a colour change**

time how long a cross under the flask takes to disappear

**the Haber process**

**hydrogen** from **methane** (natural gas); **nitrogen** from the **air**

**its conditions**

about 450 °C, 200 atm, iron catalyst – a compromise between rate and yield

**the Contact process**

sulfur dioxide to sulfur trioxide over vanadium(V) oxide, about 450 °C and 2 atm

Both use a **compromise** temperature: hotter is faster but gives less yield, because both forward reactions are exothermic.

## Exam tips

- Explain a faster rate with **collision theory**: successful collisions happen more often. Concentration, pressure and surface area all raise how often particles collide; a higher temperature does that *and* gives more particles enough energy.
- A **catalyst** lowers the activation energy and speeds the reaction up, but is not used up and does *not* change  $\Delta H$  or the position of equilibrium.
- At **equilibrium** the forward and reverse rates are equal and the concentrations stop changing – it does *not* mean the amounts of reactant and product are equal.
- Use Le Chatelier's principle: raising temperature favours the endothermic direction; raising pressure favours the side with fewer gas molecules; adding a substance shifts the equilibrium away from it.
- **Redox** – OIL RIG: Oxidation Is Loss, Reduction Is Gain of electrons. They always happen together, and an oxidising agent is itself reduced.



## Worked example – speeding up a reaction

**Powdered marble reacts with acid faster than a single lump. Explain using collision theory.**

- A reaction needs particles to **collide** with at least the **activation energy**.
- Powder has a much larger **surface area** exposed to the acid.
- So there are **more collisions per second** between acid particles and the marble.
- More successful collisions per second  $\Rightarrow$  a **faster rate**.

Surface area, concentration and pressure work by giving **more collisions**. Temperature also gives **more energetic** ones; a catalyst **lowers  $E_a$** .

## Topic 6 – key takeaways

1

### Rate

more frequent / energetic  
collisions = faster

2

### Catalyst

lowers  $E_a$ ; not used up;  $\Delta H$   
unchanged

3

### Equilibrium

forward = reverse in a closed  
system

4

### Redox

OIL RIG; oxidation & reduction  
together

## Check yourself

- 1 Why does a powder react faster than a lump?
- 2 How does a catalyst speed up a reaction?
- 3 What is equal at equilibrium?
- 4 In OIL RIG, what is oxidation?



**Answers** 1. more surface, so more collisions forward & reverse rates 2. it lowers the activation energy 3. the 4. loss of electrons