

How fast do reactions go?

Chemistry grades 9-10 (GCSE · IGCSE · O Level · Class 9-10) · Lesson

Quick answer: The rate of reaction is how quickly reactants are used up or products form, measured as amount per unit time. Collision theory says particles must collide with at least the activation energy. Higher concentration, larger surface area, higher temperature and catalysts all increase the frequency of successful collisions.

Level: Chemistry grades 9-10 (GCSE · IGCSE · O Level · Class 9-10) · **Curriculum:** HS-PS1-5

What is the rate of a reaction?

Some reactions are over in a flash, like a firework. Others take years, like rust forming on a bridge. The **rate of reaction** tells you how fast reactants are used up or products are made. A fast reaction has a high rate and finishes in a short time.

Definition

The **rate of reaction** is the amount of reactant used up, or product formed, per unit time. It is measured in cm^3/s for a gas volume, g/s for a mass, or mol/s for moles.

Formula

mean rate = amount of product formed $\div$ time taken

Use cm^3 and s for a gas volume, g and s for a mass change.

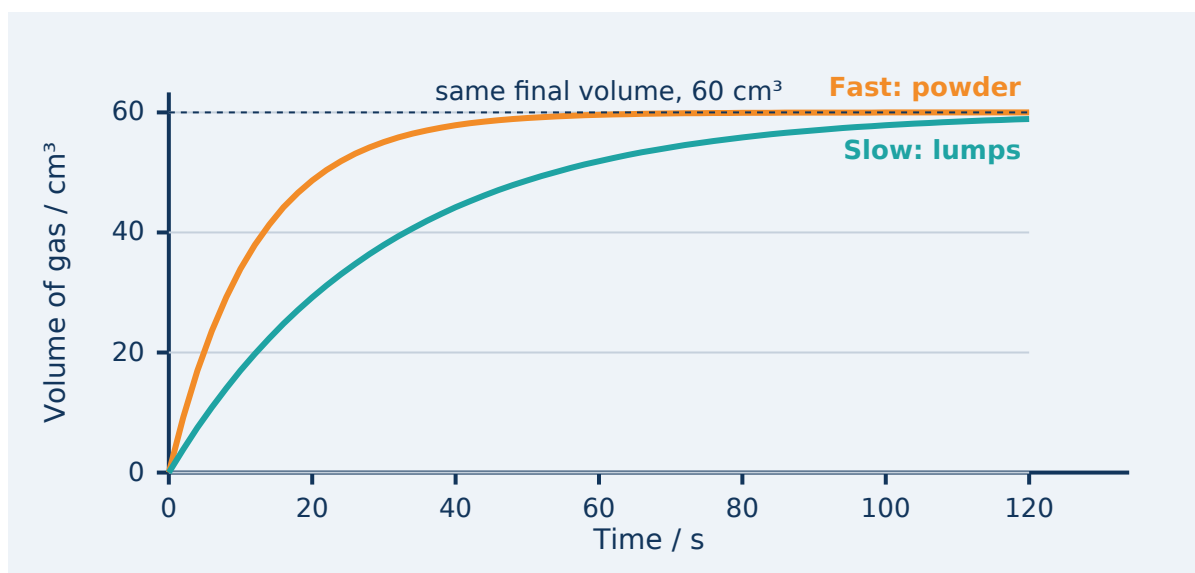
You can follow a reaction by collecting gas in a syringe, by weighing a flask that loses gas, or by timing how long a cross takes to disappear as a precipitate forms.

Worked example

Wanjiru drops marble chips into hydrochloric acid and collects 48 cm^3 of carbon dioxide in 80 s. Calculate the mean rate.

Rate = $48 \div 80 = 0.60 \text{ cm}^3/\text{s}$.

On a graph of gas volume against time, the steepness (gradient) is the rate. The curve starts steep, then flattens as the reactants are used up. When the line is flat, the reaction has stopped.



A steeper line means a faster reaction; both lines level off when a reactant is used up.

What is collision theory?

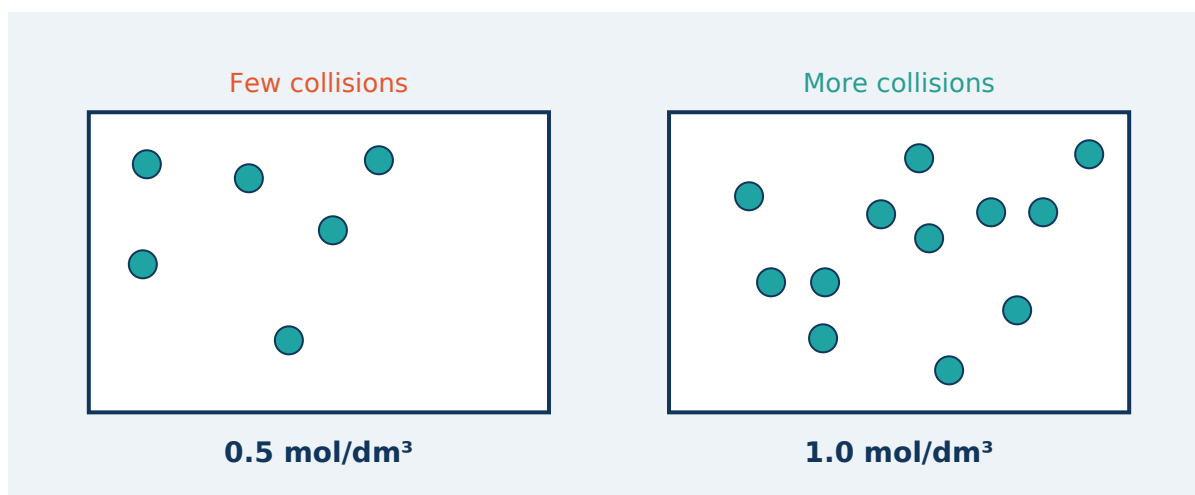
Key law

Particles can only react when they collide with at least the activation energy and in the right orientation. Any change that makes successful collisions more frequent increases the rate of reaction.

Most collisions are not successful because the particles bounce apart. Factors that raise the rate either make collisions more frequent or make a bigger share of them energetic enough.

How do concentration and pressure change the rate?

In a more concentrated solution there are more particles in each cm^3 , so they collide more often. For gases, raising the pressure squeezes the same particles into a smaller volume, which has the same effect.



Doubling the concentration puts twice as many particles in the same volume.

How does surface area change the rate?

A solid can only react at its surface. Crushing a solid into powder exposes far more surface, so more particles can be hit at the same time.

Worked example

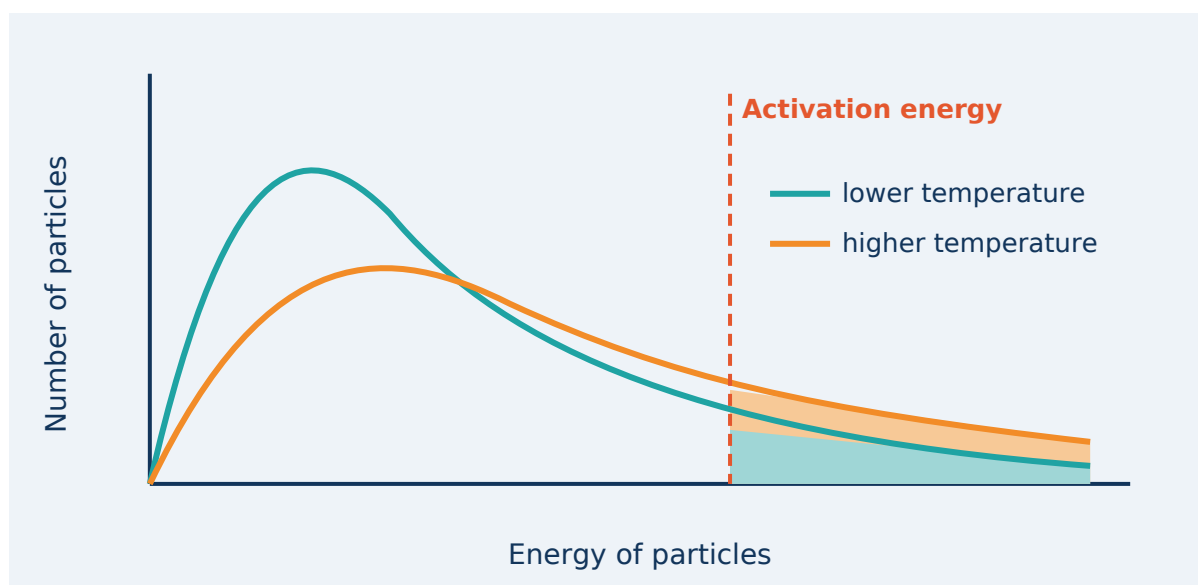
A cube of marble has sides of 2 cm. It is cut into eight cubes of 1 cm. Compare the total surface areas.

Before: $6 \times (2 \times 2) = 24 \text{ cm}^2$. After: each small cube has $6 \times 1 = 6 \text{ cm}^2$, and $8 \times 6 = 48 \text{ cm}^2$. The surface area **doubles** ($24 \rightarrow 48 \text{ cm}^2$) while the volume stays 8 cm^3 .

Change	Effect on particles	Effect on rate
Higher concentration	More particles per cm^3	More frequent collisions, faster
Larger surface area	More exposed solid particles	More frequent collisions, faster
Higher temperature	Faster, more energetic particles	More frequent and more successful collisions, faster
Catalyst	Lower activation energy	More collisions succeed, faster

How does temperature change the rate?

Heating makes particles move faster, so they collide more often. More importantly, a bigger fraction of the particles now carry at least the activation energy. Both effects raise the number of successful collisions each second.



At the higher temperature, more particles have at least the activation energy.

Worked example

Kofi finds that a reaction takes 80 s at 20°C . As a rough rule, the rate doubles for every 10°C rise. Estimate the time at 40°C .

Two rises of 10°C give $2 \times 2 = 4$ times the rate, so the time is $80 \div 4 = \mathbf{20 \text{ s}}$.

What does a catalyst do?

A **catalyst** speeds up a reaction without being used up. It gives the particles an alternative pathway with a lower activation energy, so more collisions succeed at the same temperature. Examples include manganese(IV) oxide for breaking down hydrogen peroxide, iron in the Haber process and enzymes in your body.

Common mistake

Do not say that heating only works because particles collide more often. The main reason is that many more particles reach the activation energy. Also, a catalyst does not change the amount of product, only how quickly it forms.

Lina's tip

When a question asks you to explain a rate change, use this chain: what changed, then collisions (frequency or energy), then successful collisions, then rate.

Key points to remember

- Mean rate = amount formed ÷ time (for example cm^3/s)
- A steeper line on a volume-time graph means a faster rate
- Reaction needs collisions with at least the activation energy
- Concentration, surface area and temperature raise the number of successful collisions
- A catalyst lowers the activation energy and is not used up

How fast do reactions go?: frequently asked questions

How do you calculate the rate of a reaction?

Divide the amount of product formed, or reactant used, by the time taken, for example cm^3 of gas per second. On a graph, the rate is the gradient of the line.

Why does a higher temperature speed up a reaction?

The particles move faster, so they collide more often. Far more of them also have at least the activation energy, so a larger share of the collisions lead to reaction.

Why does powdered solid react faster than lumps?

Powder has a much larger surface area, so more particles are exposed to the other reactant at the same time. This gives more collisions every second.

What does a catalyst do?

It provides a different reaction pathway with a lower activation energy. More collisions are successful, and the catalyst is not used up.

Why does the line on a gas volume graph level off?

One of the reactants has been used up, so no more gas is made. The line becomes horizontal

when the reaction has finished.