Rates of Reaction

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Rates of Reaction

Some reactions are over in a fraction of a second — an explosion, a burning match. Others take years, like the slow rusting of iron. The rate of a reaction determines how quickly reactants are converted into products, and controlling it is central to everything from food storage to industrial manufacturing.

Definition of Rate of Reaction

The **rate of reaction** is the change in concentration of a reactant or product per unit time, measured at a stated temperature.

Practically, rate can be followed by measuring any property that changes as the reaction proceeds:

- Volume of gas produced over time
- Mass lost over time (if gas escapes from an open system)
- Change in colour (if a coloured substance is formed or consumed)
- Time for a precipitate to obscure a mark beneath the flask

Collision Theory

Collision theory explains why reactions happen and why their rates change. For a reaction to occur, particles must:

- 1. **Collide** with each other.
- 2. Collide with **sufficient energy** — at least the activation energy (E_a).
- 3. Collide with the **correct orientation** for bonds to break and reform.

Only a fraction of all collisions in a given mixture meet all three conditions. These are called **successful (effective) collisions**. Anything that increases the frequency of successful collisions increases the rate.

Activation energy(E_a

) is the minimum energy required for a collision to lead to a reaction. Particles must possess at least this energy when they collide — if they do not, they simply bounce apart unchanged.

Factors Affecting Rate of Reaction

1. Concentration

Increasing concentration increases the rate.

In a more concentrated solution, there are more particles per unit volume. Particles collide more frequently, so the number of successful collisions per second rises.

> 📋 **Example**
> Marble chips (CaCO) react more vigorously and quickly in concentrated hydrochloric acid than in dilute hydrochloric acid, even if the marble chips are the same size.

For gases, increasing **pressure** has the same effect — compressing the gas brings particles closer together, increasing collision frequency.

2. Temperature

Increasing temperature increases the rate.

Two effects combine at higher temperatures:

- Particles move faster, so they collide more frequently.
- More importantly, a larger proportion of particles now possess energy greater than or equal to the activation energy — so a higher fraction of collisions are successful.

A temperature rise of 10 °C roughly doubles the rate of many reactions because of the dramatic increase in the proportion of energetic particles.

3. Surface Area

Increasing surface area increases the rate.

When a solid reactant is broken into smaller pieces or powdered, more of its particles are exposed at the surface and available to collide with the other reactant. The number of potential collisions per unit time increases even though the total amount of substance is the same.

📋 **Example**

Powdered calcium carbonate reacts much faster with hydrochloric acid than marble chips of the same mass. Both contain the same amount of $CaCO_3$ but the powder has enormously more surface area exposed.

📝 **Exam Tip**

Finely divided solids can react explosively with air if dispersed as a dust cloud — flour mills and coal mines must be carefully ventilated for this reason. The enormous surface area makes combustion almost instantaneous.

4. Catalysts

A **catalyst** is a substance that increases the rate of a reaction without being permanently chemically changed itself. It is not consumed in the reaction — the same amount of catalyst is present at the start and end.

Catalysts work by providing an **alternative reaction pathway with a lower activation energy**. More particles have enough energy to react, so more successful collisions occur per second.

Catalysts do **not** change:

- The overall energy change (ΔH) of the reaction
- The products formed
- The position of equilibrium (they help equilibrium be reached faster, not shifted)

Process	Catalyst
Decomposition of hydrogen peroxide	Manganese(IV) oxide (MnO_2)
Haber process (ammonia synthesis)	Iron
Contact process (sulfuric acid)	Vanadium(V) oxide (V_2O_5)
Catalytic cracking of petroleum	Zeolite (aluminium silicate)

Enzymes: Biological Catalysts

Enzymes are biological catalysts — proteins that speed up chemical reactions in living organisms. They work in the same way as inorganic catalysts (lowering activation energy), but they are highly specific: each enzyme catalyses only one type of reaction.

Key features of enzymes:

Feature	Detail
Structure	Proteins with a specific active site shape
Specificity	Each enzyme fits only one substrate (lock-and-key model)
Optimum temperature	Maximum activity at a specific temperature (typically 37 °C in humans)
Denaturation	At high temperatures, the enzyme's shape changes permanently — it becomes inactive
pH sensitivity	Enzymes also denature outside their optimum pH range

Measuring Rates in the Laboratory

Several practical methods can be used to monitor how quickly a reaction proceeds:

Gas collection: collect gas in a syringe or upturned burette filled with water. Record volume at regular time intervals. Plot volume against time.

Mass loss: place the reaction vessel on a balance. If gas escapes, mass decreases. Record mass at intervals.

Colour change/clock method: measure the time for a colour to appear or disappear. The reaction between sodium thiosulfate and hydrochloric acid forms a sulfur precipitate — the time for the mixture to obscure a cross drawn on paper beneath the flask is measured.

$$\mathrm{Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + SO_2(g) + S(s) + H_2O(l)}$$

Interpreting Rate Graphs

A graph of volume of gas (or mass lost) against time has a characteristic shape for most reactions:

<RateOfReactionCurve />

- The curve is **steep at the start** — reactant concentrations are highest, so collisions are most frequent.
- The gradient **decreases over time** — as reactants are used up, collisions become less frequent.
- The curve **flattens to a plateau** — the reaction is complete; no more product is formed.

A **steeper gradient** at any point means a faster rate. Comparing the initial gradients of two curves shows which set of conditions produced a faster reaction.

Changes in Conditions on Rate Graphs

Change	Effect on curve
Higher concentration / finer powder / higher temperature / catalyst	Steeper initial gradient; plateau reached sooner
More reactant (but same rate)	Same initial gradient; plateau reached at higher value

Rates and Equilibrium

Some reactions are **reversible** — the products can reform the reactants. In a closed system, a **dynamic equilibrium** is established when the forward and reverse reactions occur at equal rates. At this point, concentrations no longer change even though both reactions continue. Altering temperature, pressure, or concentration shifts the equilibrium position; a catalyst speeds up both directions equally and does not change where equilibrium lies.