1: What is rate of reaction? And why is it important to know?

Rate means speed. We can increase the rate of a reaction to make a product faster. Scientists want to increase the speed of a reaction but also **reduce the energy** needed (£££).

2: Collision Theory – describing how reactions happen

For a reaction to take place particles must:
1. collide in the **right orientation** (direction) to react
2. collide with **enough energy** to react.

frequency of collisions: how often particles bump into each other
Higher frequency is likely to lead to a faster reaction so long as the particles have enough energy when they collide.

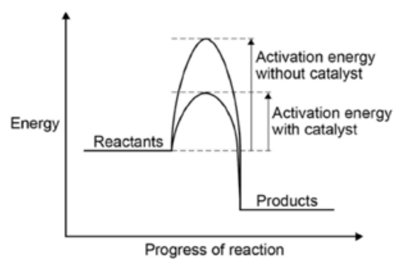
activation Energy (E_A): minimum amount of energy particles must have to react.

The effect on the rate of increasing a variable:

Variable to increase	Effect on collisions	Effect on collision success	Overall rate change
surface area	more frequent	no change	increase
concentration	more frequent	no change	increase
pressure	more frequent	no change	increase
temperature	more frequent	more energy, more collisions have E_A or higher	increase
use a catalyst	no change	reduces E_A so more collisions have E_A or higher	increase

Reaction profiles show the relative difference in the energy level of the reactants compared to the products.

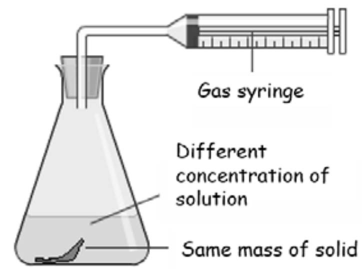
You also learnt about reaction profiles in C7 Energy Changes.



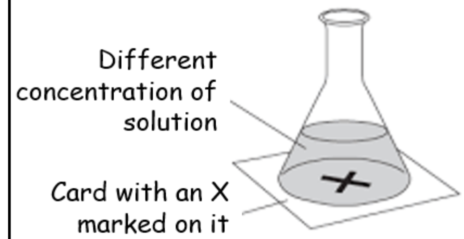
3: Measuring rate as concentration changes (required practical)

As the concentration of a reagent is increased, the reaction rate increases. This can be observed and tested for in two ways:

Measure volume of gas produced:



Measure the turbidity of solution (how quickly the X disappears):



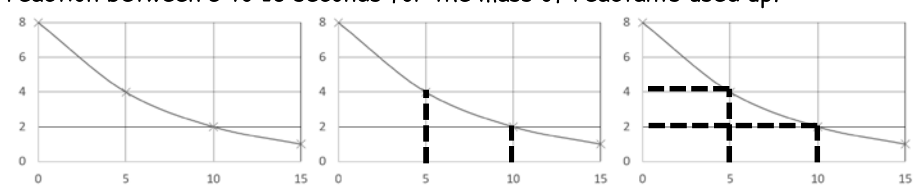
4: Calculating the rate of a reaction

Mean rate of reaction =
quantity of reactants used (g) ÷ time taken (s) = rate (g/s)
or
volume of gas produced (cm^3) ÷ time taken (s) = rate (cm^3/s)

Tip: always calculate rates per second (1 min = 60 seconds)

(HIGHER ONLY) need to know **mol/s** as well!

Steps to calculating the mean rate from a graph: What is the mean rate of reaction between 5 to 10 seconds for the mass of reactants used up?



Mean rate of reaction = quantity of reactants used (g) ÷ time taken (s)
= (4 g – 2 g) ÷ (10 s – 5 s)
= 0.5 g / s

A **catalyst** is a chemical added to a reaction which increase the rate.

Catalysts are **not used up** and be used over and over many times.

Enzymes are biological catalysts.

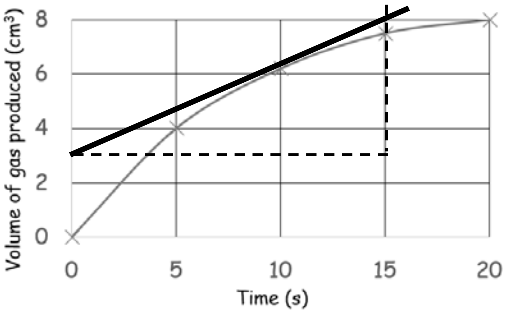
Catalysts **are not reactants**, so they are **not written in the chemical equation.**

They increase the rate by providing an **alternative reaction pathway** with **lower activation energy.**

Different reactions need different catalysts.

Drawing a tangent

Tangent: a straight line that touches the curve.
E.g. “draw a tangent at 10 seconds” (bold line)



(HIGHER ONLY)
Rate of reaction **AT** a particular time.
Draw a tangent **AT** the time and then calculate its gradient (See below)

Calculating the gradient of a tangent (HIGHER ONLY)

With the tangent draw, make the biggest right angled triangle you can and calculate “rise over run” dotted lines, answer is:
Rise (8-3) ÷ Run (15 – 0) = 5 ÷ 15
The rate (gradient) = 0.33 cm^3/s

6: **Reversible reactions** have a special arrow: $\rightleftharpoons$

Some reactions are **reversible**: the products can react and turn back into the starting reactants.

The direction of reversible reactions can be changed by changing the **conditions**.

$$\text{ammonium chloride} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{ammonia + hydrogen chloride}$$

In the example below, adding heat will form the white solid, cooling down will form the blue solid.

Forward reaction

$$\text{Hydrated copper sulfate (blue)} \underset{\text{(exothermic)}}{\overset{\text{(endothermic)}}{\rightleftharpoons}} \text{copper sulfate + water (white)}$$

Backward reaction

7: **Equilibrium:** equal in both directions

Closed system: when reactants and products can t escape (like a flask with a lid on)

In a **closed system** a **reversible reaction** will reach **equilibrium**.

Dynamic equilibrium is when the **forward reaction** is happening at the **same time** and the **same rate** as the **backward reaction**.

Note: This **does not** mean that there is 50% product and 50% reactants. Just that the quantities of reactants and products doesn t change overall.

Le Chatelier’s Principle (HIGHER ONLY):

“If a change is made to the conditions of a system that is at equilibrium, then the system will respond to counteract the change”.

system: the reaction
counteract: reverse the change
equilibrium: equal in both directions

We can use le Chatelier's principle to predict the effect of changing conditions in a system such as:

- warming up or cooling down a reaction
- adding more of some of the chemicals involved
- removing some of the chemicals involved

If the concentration of a reactant is increased, more products will be formed until equilibrium is reached again.

If the concentration of a product is decreased, more reactants will react until equilibrium is reached again.

Example:

$$N_2O_{4\,(g)} \underset{\text{(endothermic)}}{\overset{\text{(exothermic)}}{\rightleftharpoons}} N_{2\,(g)} + 2O_{2\,(g)}$$

In this example, if you add heat to the equilibrium, it will move in the endothermic direction to attempt to lose the heat.

For reactions between gases, if you increase **pressure**, the equilibrium position will shift to favour the side with fewer moles of gas (as shown by the symbol equation) because every gas occupies the same volume of space. *In the example above, there is 1 mole of N_2O_4 for every N_2 and $2O_2$. So, increasing pressure will favour N_2O_4 (fewer moles), whereas decreasing pressure will favour N_2 and $2O_2$ (more moles).*