Summary Sheet Year 9 Rate of Reaction

Measuring Rate

Remember that reactants → products

The rate of a chemical reaction can be found by measuring the amount of a reactant used or the amount of product formed over time.

$$\text{rate} = \frac{\text{change (in cm}^3\text{ or g)}}{\text{time (s)}}$$

The amount can be measured by the change in mass in grams or by a volume in cm^3.

Units of rate of reaction may be given as **g/s** (change in mass divided by time) or **cm^3/s** (change in volume divided by time)

Eg.1 If $25cm^3$ of gas is made in 10 seconds, rate = $25cm^3$/10s = **2.5 cm^3/s**

Eg.2 If 7g of metal dissolves in 1 minute 20 seconds (convert to 80 seconds), rate = 7g/80s = **0.0875 g/s**

Interpreting Graphs

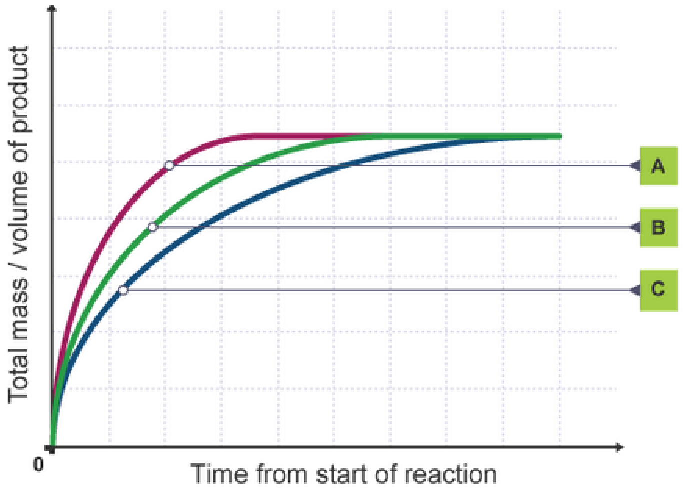
Mean (average) rate of a reaction can be calculated from supplied data or from graphs - the **steeper the gradient** on the graph, the **faster the rate** of reaction

eg. if a reaction makes 40 cm^3 gas in 60 seconds

mean rate = 40 cm^3 ÷ 60 s = **0.67 cm^3/s**

From the graphs shown here **reaction A is faster** as the curve (and its gradient) is **steeper**.

There is a greater increase in the amount of product over a shorter period of time.



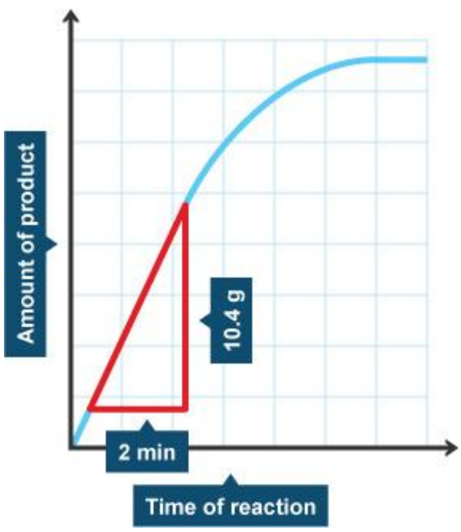
Extension – Using gradients to find rate

Can you calculate the gradient of a tangent to the curve on these graphs as a measure of rate of reaction at a specific time?

Remember that 2 minutes = 120 seconds

Gradient here would be:

10.4 g ÷ 120 s = **0.087 g/s**



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Collision Theory

Rate of reaction can be explained using the idea of **particles** and **collisions**.

For a reaction to take place:

- particles must collide
- and they must collide with the required **ACTIVATION ENERGY (E_a)**

The activation energy, E_a, is the **minimum energy** needed for a reaction to take place.
The **more frequent collisions** between particles taking place, the **faster** the reaction.

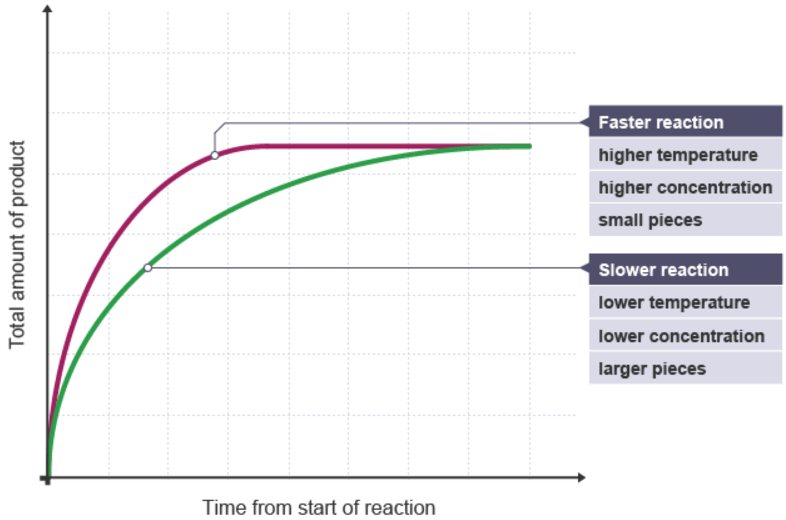
Reactions are always fastest at the beginning, when there are more reactant particles present, so more frequent collisions occur.

Reactions always slow down as particles get used up.

FACTORS AFFECTING RATE (always explain using collision theory)

1. **Concentration**
   - more concentrated
   - more particles in the same volume
   - more frequent collisions
   - Faster rate of reaction

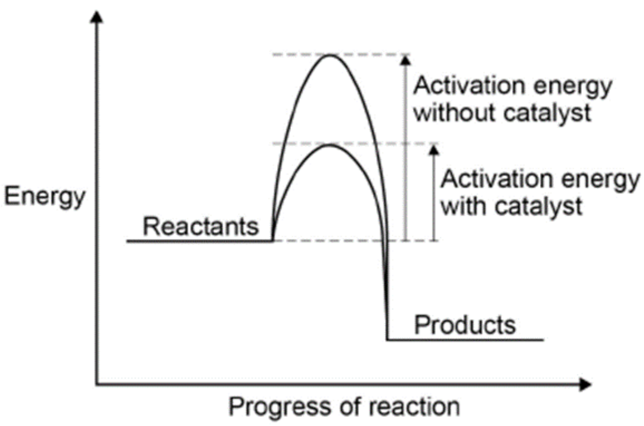
2. **Pressure**
   - More particles in the same volume
   - particles closer together
   - more frequent collisions
   - Faster rate of reaction



3. **Temperature**
   - higher temp
   - particles have more energy so move faster
   - more frequent collisions (and harder ones with more energy - more likely to have the activation energy, E_a)
   - Faster rate of reaction

4. **Surface Area**
   - smaller pieces
   - more surface area so more exposed particles
   - more frequent collisions
   - Faster rate of reaction

5. **Catalyst**
   - lowers activation energy, E_a
   - more collisions now have the activation energy
   - Faster rate of reaction



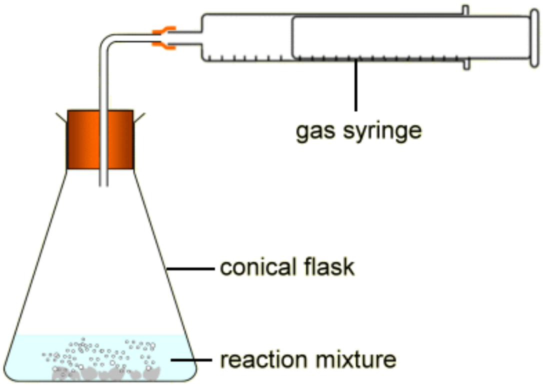
Summary Sheet Year 9 Rate of Reaction

Rate of Reaction Practical Work

1) Measure how quickly a gas is being produced
2) Time how quickly a reaction gets to a fixed point (colour change)

Method 1 eg. magnesium reacting with hydrochloric acid to make magnesium chloride and hydrogen, H_2

$Mg_{(s)} + 2\,HCl_{(aq)} \longrightarrow MgCl_{2\,(aq)} + H_{2\,(g)}$



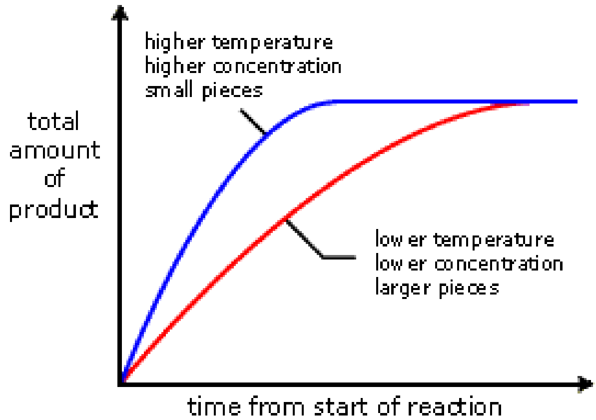
The **gas syringe** measures the volume of gas produced accurately.

A **stopwatch** would also be needed to time the reaction.

The volume of gas is recorded at regular timed intervals, eg. every 10 seconds.

To investigate the **effect of concentration of acid in method 1**

- Independent variable = concentration of acid (what **I** changed)
- Dependent variable = volume of gas (what was measured)
- Control variables (what we keep the same) would be:
  - volume of acid (use a measuring cylinder)
  - mass of magnesium (use a balance)
  - surface area of magnesium
  - temperature



The results might be used to plot a graph.

The **steeper** the graph, the **faster** the reaction.

It gradually **slows down as particles get used up**.

This means **less frequent collisions** are happening.

The reaction **stops** when one of the **reactants** (the limiting factor) gets **used up** and the **graph goes flat**.

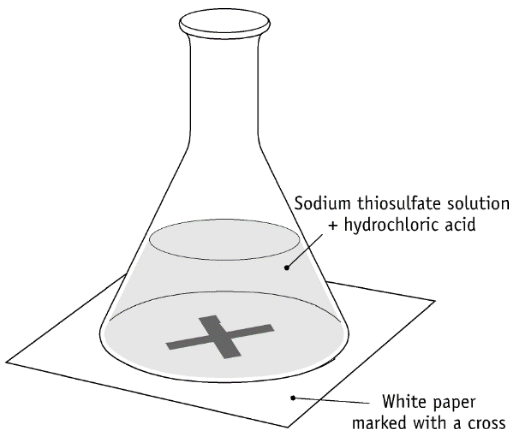
Method 2 eg. a reaction between sodium thiosulfate and hydrochloric acid, that produces a cloudy precipitate of sulfur (S)

$Na_2S_2O_{3\,(aq)} + 2\,HCl_{(aq)} \longrightarrow 2\,NaCl_{(aq)} + S_{(s)} + SO_{2\,(g)} + H_2O_{(l)}$

Reaction is timed from when the chemicals are mixed, until the point at which the cross is no longer visible below the conical flask.

The cross eventually disappears as enough **sulfur** is formed, which is a **solid** so blocks out the cross.

The **faster the reaction, the less time it takes**.



If you were investigating the **effect of temperature using method 2,** could you list the independent, dependent and 3 or 4 control variables?