1 – Collision theory

Collision theory	For a reaction to occur, particles must **collide** with **sufficient energy**.
Activation energy	**Minimum** amount of **energy** that reactant particles must have when they **collide** in order to **react**.

2 – Factors affecting rate of reaction

Temperature	**Increase** temperature → particles have **more energy** so move **faster** → collide more **frequently** → more collisions have **activation** energy → greater chance of **successful collisions**.	
Concentration/ Pressure	**Increase concentration** of solution → more particles in same **volume** → particles collide more **frequently**.	**Increase pressure** of gas → same number of particles **occupy smaller** space → particles are **closer** → collide more **frequently.**
Surface area	Breaking a solid into **smaller** pieces **increases surface area** to **volume ratio** → particles have more **area** to **collide** with → **collisions** occur more **frequently**.	
Catalyst	A substance that **speeds up rate** of reaction **without being used** up in the reaction. They work by **decreasing** the **activation** energy needed by providing an **alternative** reaction **pathway**.	

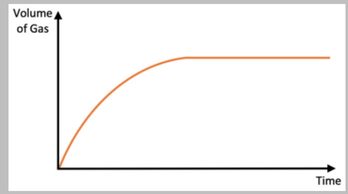
3 – Measuring rates of reaction

Calculation	$mean\ rate = \frac{quantity\ of\ reactant\ used}{time\ taken}$ units = g/s $mean\ rate = \frac{quantity\ of\ product\ formed}{time\ taken}$ units = g/s or cm^3/s
Measuring mass	Measure mass at the **start** and **end** of a reaction to measure **mass** of **gas** produced. Not suitable for **hydrogen** due to small mass.
Measuring volume	**Volume** of **gas produced** can be measured using a **gas syringe** or an **upside-down measuring cylinder** filled with water.
Precipitation	Can be used when **initial solution** is **transparent,** and the **product** is a **precipitate,** so the solution becomes **opaque**. The **cloudiness** of the solution is called the **'turbidity'**. Observe a **mark** through the solution and measure **time** taken to **disappear**.

4 – Rate of reaction graphs

Shape of graphs

The **steeper** the line, the **faster** the rate. As the line gets **less steep**, the rate is **slowing** down. A **horizontal** line means the reaction is **finished**.



Calculating mean rate

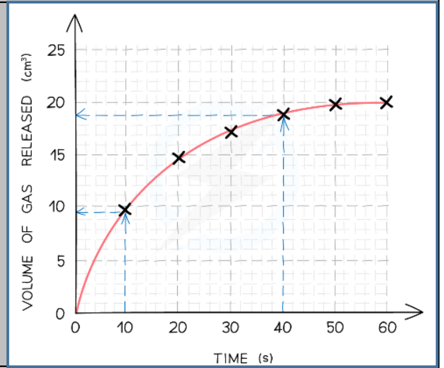
$mean\ rate = \frac{change\ in\ y}{change\ in\ x}$

e.g

y = 19 – 8.5 = 10.5

x = 40 – 10 = 30

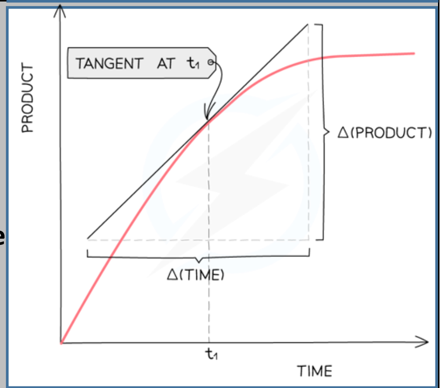
$10.5/30 = 0.35\ cm^3/s$



Calculating instantaneous rate

A **tangent** need to be drawn to calculate **change** in **x and y** then the rate of reaction at a **particular point**.

1. Draw a **tangent** at the **time** the **rate** needs to be **calculated**.
2. Calculate the **gradient** of the tangent: **change in product ÷ change in time**
3. Gradient = **rate of reaction.**



GCSE Science

Chemistry C6 – Rate and extent of chemical change

5 – Required practical methods	
Observing colour change	1. 10cm^3 of **sodium thiosulfate** and 40cm^3 of **water** into **conical flask** (diluted to 8g/dm^3). 2. Place flask on a **white tile** with **black cross**. 3. Add 10cm^3 of dilute **HCl** and start **timer**. 4. Gently **swirl** the flask and stop timer when the **cross** is **no longer visible**. 5. **Repeat** using **different concentrations** of sodium thiosulfate. 6. **Repeat** whole investigation twice more and **calculate** a **mean** for each concentration.
Measuring volume of gas produced	1. Add 50cm^3 of **1.0mol/dm^3 HCl** to a **conical flask**. 2. Fit a **bung** and **delivery tube** to the conical flask. 3. Half fill a **trough** with **water**. 4. Fill a **measuring** cylinder with **water**, ensure it stays full and **invert** it into the **water** trough, with the **delivery tube** positioned into it. 5. Add a **3cm** strip of **magnesium** to the **conical** flask and **replace** the bung, then start the **timer**. 6. Record the **volume** of hydrogen **gas** given off in **10 second** intervals. 7. **Repeat** using **1.5mol/dm^3** acid.

6 - Reversible reactions and equilibrium		
Reversible reaction	A reaction where the **products** can **react** to produce the **original reactants**. Shown by: $\rightleftharpoons$	
Equilibrium	The point at which the **rates** of the **forward** and **backward** reactions in a **reversible** reaction are the **same**. The amounts of **reactants** and **products** in the reaction **container don't change**.	
Positions of equilibrium	If the **equilibrium** lies to the **right:** the **concentration** of **products** is **higher** than the **reactants**.	If the **equilibrium** lies to the **left**: the **concentration** of **reactants** is **greater** than the **products**.
Changing the direction	**Heating** a **reversible** reaction moves it in the **endothermic direction**. **Cooling** a **reversible** reaction moves it in the **exothermic direction**.	

7 – Le Chatelier's principle (HT)	
Le Chatelier's principle	If you **change** the **conditions** of a **reversible** reaction, at **equilibrium**, the system will try and **counteract** the **change**.
Changing temperature	**Decrease** temperature → reaction moves in **exothermic direction**. **Increase** temperature → reaction moves in **endothermic direction**.
Changing concentration	**Increase concentration** of **reactants** → system makes more products (the **forward reaction increases**). **Decrease concentration** of **products** → system tries to make more products by reducing reactants (**forward reaction increases**).
Changing pressure	Only affects **equilibrium** involving **gases**. **Increase** pressure → equilibrium tries to reduce it and moves to the side with **fewer molecules**. **Decrease** pressure → equilibrium tries to increase it and moves to the side with **more molecules**.

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Chemistry C6 – Rate and extent of chemical reactions.