

## 4.6 Rate and Extent of Chemical Change

### Rates of Reaction

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:

$$\text{Rate of reaction} = \frac{\text{Amount of reactant used}}{\text{Time}}$$

$$\text{Rate of reaction} = \frac{\text{Amount of product formed}}{\text{Time}}$$

The quantity of reactant or product can be measured by the mass in grams or by a volume in  $\text{cm}^3$ .

The units of rate of reaction may be given as  $\text{g/s}$  or  $\text{cm}^3/\text{s}$ .

The quantity of reactants can also be measured in terms of moles and for rate of reaction the unit can be measured in  $\text{mol/s}$ .

Be able to calculate the **mean rate of a reaction** from given information about the quantity of a reactant used or the quantity of a product formed and the time taken. This data could be taken from any two points on a graph of concentration against time

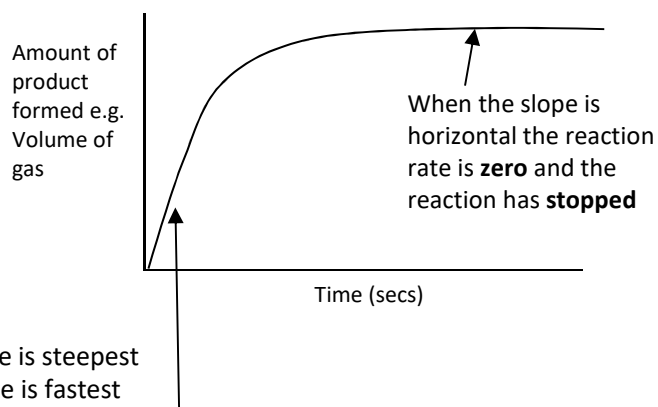
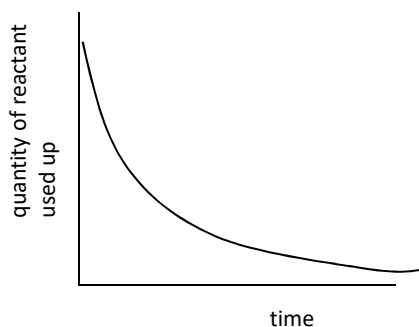
#### Example

If  $25\text{cm}^3$  of hydrogen gas is produced in 10 seconds then the rate of reaction is  $25/10 = 2.5\text{cm}^3/\text{s}$

### Using graphs to show reaction rates

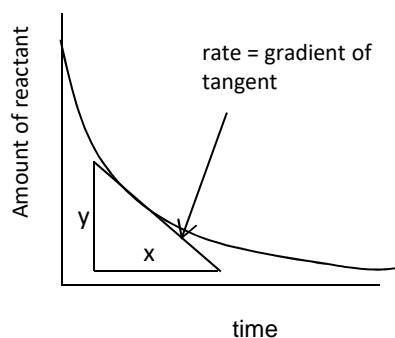
draw, and interpret, graphs showing the quantity of product formed or quantity of reactant used up against time

The **slope/gradient** of these graphs is a measure of reaction rate.



Where the slope is steepest the reaction rate is fastest

### Calculating the gradient



Draw tangents to the curves on these graphs. The tangent can be drawn at any time on the graph. (What time does the question ask for the rate?) Calculate the gradient of the tangent by drawing a tangent triangle.

$$\text{Gradient} = \frac{\text{difference in } y}{\text{difference in } x}$$

Use the slope/gradient of the tangent as a measure of the rate of reaction at that time.

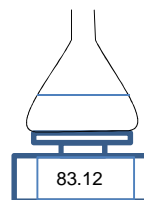
## Common ways of measuring rate:

You should be able to investigate factors which affect the rate of chemical reactions by measuring:

- the loss in mass of reactants
- the volume of gas produced
- the time for a solution to become opaque or coloured.

**Measuring mass loss.** This can be done if a heavy gas like carbon dioxide is given off. The mass drops as the gas is given off. (Hydrogen gas is too light for this method to be used)

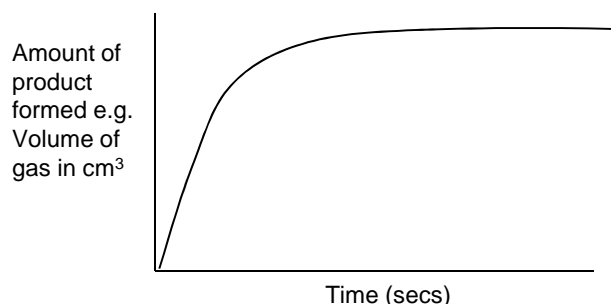
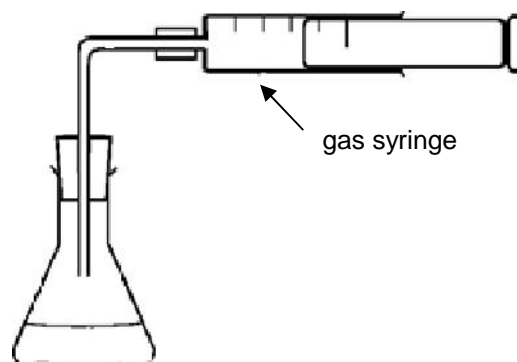
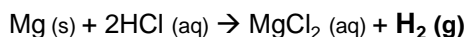
Typically used for marble chips ( $\text{CaCO}_3$ ) and acid reaction



**Measuring the volume of gas** evolved over time using a gas syringe. Gas volume readings are taken at regular time intervals and a graph plotted.

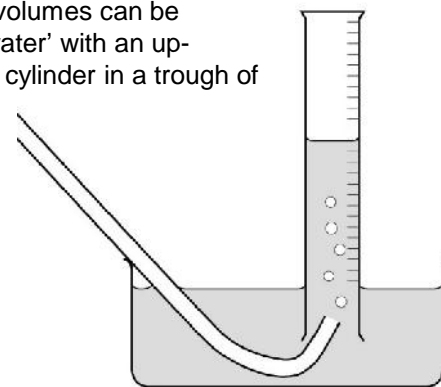
The gradient of this graph will be equal to the reaction rate. The steeper the graph the faster the reaction.

Typically used for magnesium and acid reaction



The gradient of the slope can be used as a measure of reaction rate

Alternatively gas volumes can be measured 'over water' with an up-turned measuring cylinder in a trough of water



Typical method used for marble chips ( $\text{CaCO}_3$ ) and acid (same method for magnesium).

1. Weight 10 g of marble chips on a mass balance
2. Place marble chips into a flask.
3. Measure 50 cm<sup>3</sup> of hydrochloric acid with a measuring cylinder
4. Pour hydrochloric acid into the flask, connect the gas syringe and start a timer.
5. Record the volume of gas produced every 10 seconds.



## Factors which affect the rates of chemical reactions

include:

- the concentrations of reactants in solution,
- the pressure of reacting gases,
- the surface area of solid reactants,
- the temperature
- the presence of catalysts.

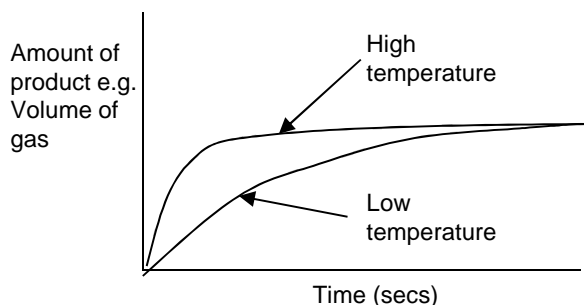
## Collision Theory

Collision theory explains how various factors affect rates of reactions. According to this theory, chemical reactions can occur only when reacting particles **collide** with each other and **with sufficient energy**.

The **minimum amount of energy** that particles must have to react is called the **activation energy**.

### Increasing Temperature

**Increasing the temperature** means the **particles gain energy** increases the **speed of the reacting** particles so that they **collide more frequently** and **more energetically**. More of the collisions are **successful** as more have **energy greater** than the **activation energy**. This increases the rate of reaction.



### Increasing Concentration

Increasing the concentration of reactants in solution, increases the number of **particles in the same volume** and so **increases the frequency of collisions** and therefore increases the rate of reaction.

### Increasing the Pressure of a Gas

Increasing the pressure of reacting gases, is the same as increasing concentration. It increases the number of gas **molecules in the same volume** and so **increases the frequency of collisions** and therefore increases the rate of reaction.

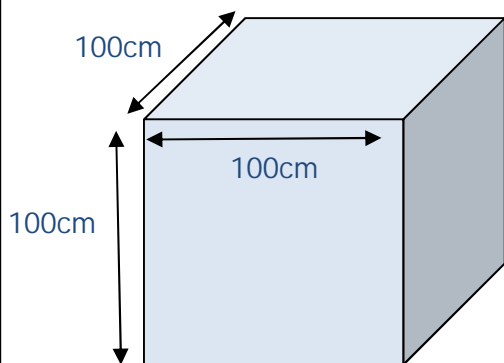
### Increasing the Surface Area

If solid reactants are in smaller pieces they have a greater surface area.

Increasing the surface area of solid reactants increases the **frequency of collisions** and so increases the rate of reaction.

be able to explain the effects on rates of reaction of changes in the size of pieces of a reacting solid in terms of surface area to volume ratio.

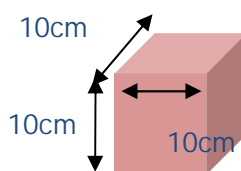
The smaller the piece the larger the surface area to volume ratio



$$\text{Volume} = 100 \times 100 \times 100 \\ = 1000\,000 \text{ cm}^3$$

$$\text{surface area} = 100 \times 100 \times 6 \\ = 60\,000 \text{ cm}^2$$

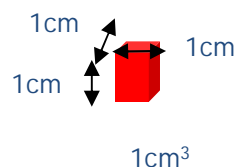
Surface area to volume ratio  
0.06:1



$$\text{Volume} = 10 \times 10 \times 10 \\ = 1000 \text{ cm}^3$$

$$\text{surface area} = 10 \times 10 \times 6 \\ = 600 \text{ cm}^2$$

Surface area to volume ratio  
0.6:1



$$\text{Volume} = 1 \times 1 \times 1 \\ = 1 \text{ cm}^3$$

$$\text{surface area} = 1 \times 1 \times 6 \\ = 6 \text{ cm}^2$$

Surface area to volume ratio  
6:1

## Catalysts

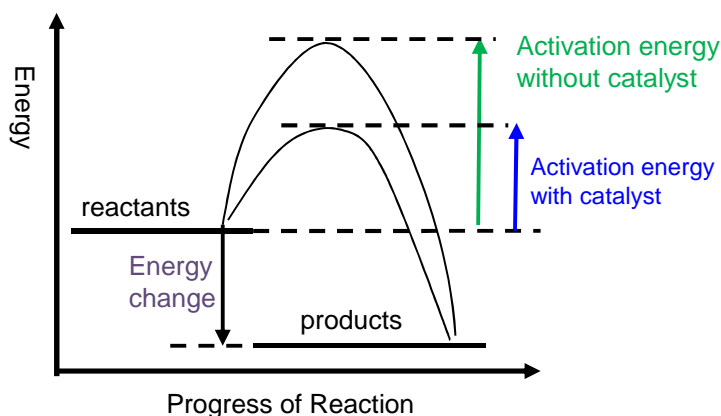
Catalysts change the rate of chemical reactions but are **not used up** during the reaction

Different reactions need different catalysts. Enzymes act as catalysts in biological systems.

Catalysts increase the rate of reaction by providing a different pathway for the reaction that has a lower activation energy.

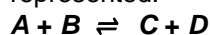
Be able to identify catalysts in reactions from their effect on the rate of reaction and because they are **not included** in the **chemical equation** for the reaction

A reaction profile for a catalysed reaction can be drawn in the following form:



## 4.6 Reversible Reactions

In some chemical reactions, the products of the reaction can react to produce the original reactants. Such reactions are called **reversible reactions** and are represented:



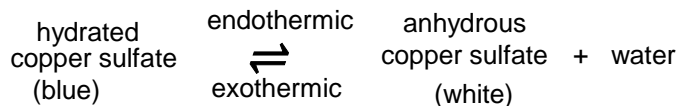
When a reversible reaction occurs in apparatus which prevents the escape of reactants and products (called a **closed system**) an **equilibrium** is reached.

**Equilibrium** is reached:

- when the **forward** and **reverse** reactions occur at **exactly the same rate**.
- When reaction occurs in apparatus which prevents the escape of reactants and products.

**At equilibrium** the amounts of reactants and products **remain constant**

If a reversible reaction is exothermic in one direction, it is endothermic in the opposite direction. The same amount of energy is transferred in each case. For example:



## Changing Conditions

The relative amounts of all the reactants and products at equilibrium depend on the conditions of the reaction.

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

For example: 
$$\text{Ammonium chloride} \xrightleftharpoons[\text{cool}]{\text{heat}} \text{ammonia} + \text{hydrogen chloride}$$

The effects of changing conditions on a system at equilibrium can be predicted using **Le Chatelier's Principle**:

If a system is at equilibrium and a **change** is made to any of the **conditions**, then the system responds to **counteract the change**.

## Changing Concentration

If the concentration of one of the reactants or products is changed, the system is no longer at equilibrium and the concentrations of all the substances will change **to counteract the change** until equilibrium is reached again.

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

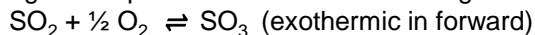
If the concentration of a **product is decreased**, **more reactants will react** until equilibrium is reached again.

## Changing Temperature

If the **temperature is increased** the equilibrium will always move in the **endothermic direction** to so that the increase in temperature is reduced.

If the temperature is increased, the yield from the endothermic reaction increases and the yield from the exothermic reaction decreases.

e.g. If temp is increased in the following reaction



The yield of  $\text{SO}_3$  will decrease as the reaction will move in the **reverse** endothermic reaction.

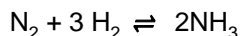
If the temperature of a system at equilibrium is increased:

- the relative amount of products at equilibrium increases for an endothermic reaction
- the relative amount of products at equilibrium decreases for an exothermic reaction.

## Changing Pressure

In gaseous reactions, if **pressure is increased** the equilibrium will move towards the side of the reaction that has the **least number of gaseous molecules** so that the increase in pressure is reduced

e.g. If pressure is increased in the following reaction



The yield of  $\text{NH}_3$  will increase as the reaction will move to the **right side** which has **fewer moles of gas**.

An increase in pressure causes the equilibrium position to shift towards the side with the smaller number of molecules as shown by the symbol equation for that reaction.

A decrease in pressure causes the equilibrium position to shift towards the side with the larger number of molecules as shown by the symbol equation for that reaction.

If a reaction has the **same number of gaseous** molecules on both sides of the equation, then increasing pressure will have **no effect** on the position of equilibrium e.g.  $2 \text{HI} \rightleftharpoons \text{H}_2 + \text{I}_2$

Increasing temperature and pressure will also increase the rate of reaction. Explain this by using the collision theory from earlier in this chapter.

Catalysts do not affect the position of equilibrium as they speed up the forward and the backward reactions by the same amount.