FACTFILE: GCSE CHEMISTRY: UNIT 2.3

Rewarding Learning

Rates of Reaction

Learning outcomes

Students should be able to:

2.3.1 demonstrate knowledge and understanding that the rate of a reaction may be determined by measuring the loss of a reactant or gain of a product over time and use the equation:

$$\text{Rate} = \frac{1}{\text{time}}$$

2.3.2 suggest appropriate practical methods to measure the rate of a reaction and collect reliable data (methods limited to measuring a change in mass, gas volume or formation of a precipitate against time) for the reaction of:

- metals with dilute acid;
- calcium carbonate/marble chips with dilute hydrochloric acid;
- catalytic decomposition of hydrogen peroxide; and
- sodium thiosulfate and acid (equation not required);

2.3.3 interpret experimental data quantitatively, for example drawing and interpreting appropriate graphs to determine the rate of reaction;

2.3.4 **describe and explain the effects on rates of reaction when there are changes in:**

- **temperature;**
- **concentration;**
- **frequency and energy of collisions between particles; and**
- **changes in particle size in terms of surface area to volume ratio (other practical activity)**; and

2.3.5 demonstrate knowledge and understanding that a catalyst is a substance which increases the rate of a reaction without being used up and recall that transition metals and their compounds are often used as catalysts.

2.3.6 explain catalytic action in terms of providing an alternative reaction pathway of lower activation energy;

Rates of Reaction

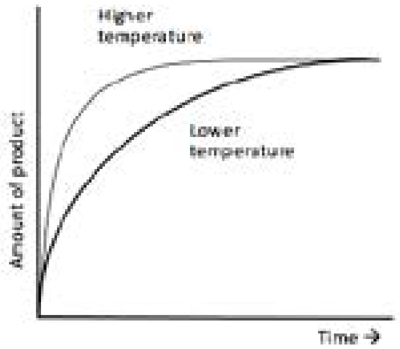
In everyday life, different reactions take place at different rates. Some reactions are slow, some are fast.

The **rate** of reaction is how quickly the reaction happens. In order for a reaction to take place, there must be a **collision** between the particles concerned. Collision theory states that in order for a reaction to take place, and to get a **successful collision**, the collision must have enough energy otherwise the particles simply bounce apart. It is economically very useful to increase the rate of a reaction in industry as this helps to reduce energy costs.

There are a number of factors which can be altered to increase or decrease the rate of a reaction. These are:

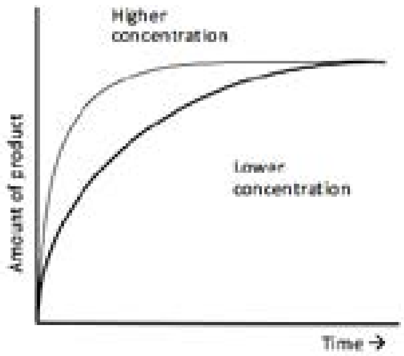
- changing the temperature
- changing the concentration
- changing the particle size or
- by adding a catalyst

Temperature: When you increase the temperature, the particles move quicker because they have more energy – they are more likely to have successful collisions and this increases the rate of the reaction. When you decrease the temperature, this slows down the movement of the particles and hence decreases the chances of a collision.



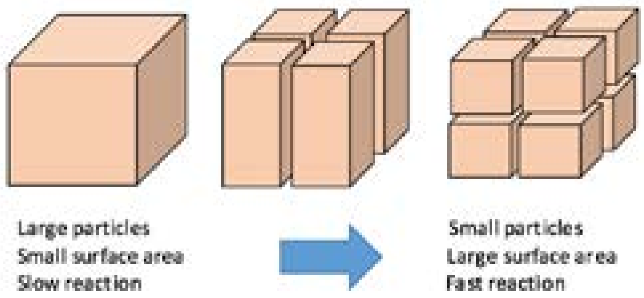
In everyday life, food spoils because of chemical reactions that occur. Food spoiling chemical reaction is slower at low temperatures as particles are moving more slowly so food remains usable for so much longer if it is stored in a freezer.

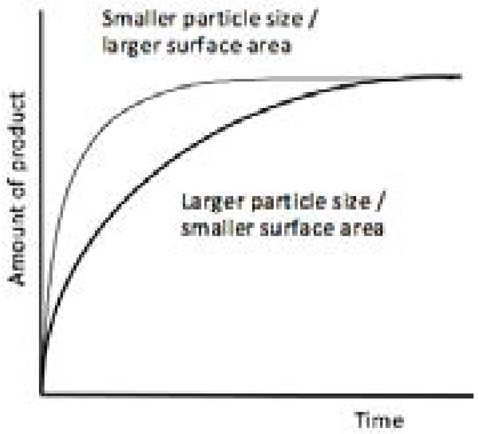
Concentration: The more particles there are present in a reaction vessel, the greater the chance of collisions. For example, increasing the concentration of an acid (whilst maintaining the overall volume of acid used) increases the number of reactant particles present and therefore increases the number of successful collisions in a given period of time. The same can be said for gaseous reactants – increasing the pressure brings the particles closer together, in effect increasing their concentration therefore increasing the number of successful collisions to occur.



Particle size / Surface Area:

If we grind a solid into a powder, we increase the surface area. If we increase the surface area, by decreasing the particle size, we increase the number of collisions which increases the number of successful collisions in a given period of time, and so increase the rate of reaction.



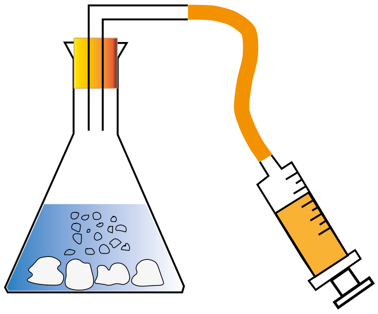


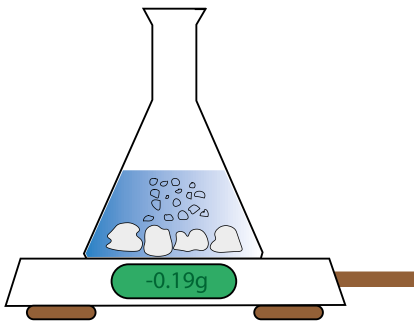
Catalysts: A catalyst speeds up a chemical reaction without being used up in the process. A good example is the production of ammonia. At room temperature, hydrogen and nitrogen do not react to make ammonia. However, if the mixture is heated to 450°C and passed through pressure of 250 atmospheres, they turn into ammonia. Iron is added as a catalyst to speed up the reaction. Scientists often use transition metals to catalyse reactions.

Measuring the Rate of a Reaction

There are two common ways to measure the rate of a reaction:

1. Measure how fast the products are formed or;
2. Measure how fast the reactants are used up.





The rate of a reaction can be calculated using the equation below:

$$\text{Rate of a reaction} = \frac{1}{\text{time}}$$

Let's look at the reaction between magnesium and dilute sulfuric acid producing hydrogen gas and magnesium sulfate.

magnesium + sulfuric acid → magnesium sulfate + hydrogen

$$Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$$

As the reaction progresses, the rate of reaction can be measured in terms of:

a) The amount of reactant used up in a given time

This could be the amount of magnesium or sulfuric acid used up per minute or;

b) The amount of product obtained in a given time

This could be the amount of magnesium sulfate or hydrogen formed per minute.

If 0.1 g of magnesium were added to dilute sulfuric acid and it took 20 seconds for the magnesium to react completely and give 100 cm^3 of hydrogen gas, the rate of reaction would be:

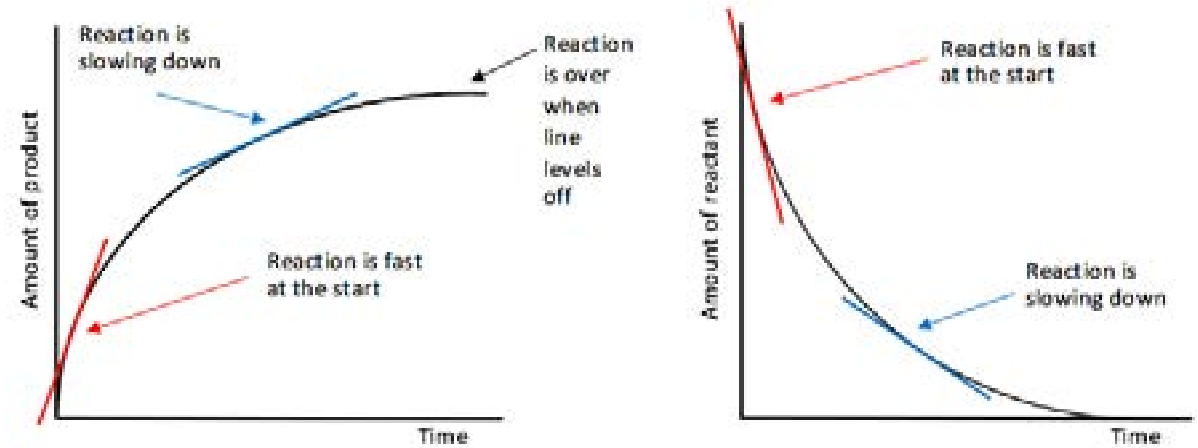
$$\text{Rate} = \frac{\text{change in mass of magnesium}}{\text{time taken}}$$

$$\text{or } \frac{\text{change in volume of hydrogen gas given off}}{\text{time taken}}$$

$= \frac{0.1}{20}$ =0.005 grams per second (g s^{-1}) or $\frac{100}{20}$ = 5 cm^3 per second (cm^3 s^{-1})

The rate calculated above is the **average rate** over 20 seconds for all of the magnesium to react or for the total volume of hydrogen to form.

The rate of a reaction can also be calculated from a graph by using it to determine the gradient.



The gradient, in the early stages of the reaction illustrated by the graph above, is very steep, demonstrating that the reaction is very fast at the beginning. Over time, the gradient becomes less steep which indicates that the reaction is slowing down because there are less remaining unreacted particles present. The curve will then level off, indicating that the reaction is over.

The following four reactions are very important:

1. Reaction of metals with dilute acid:

metal + acid ⟶ metal salt + hydrogen

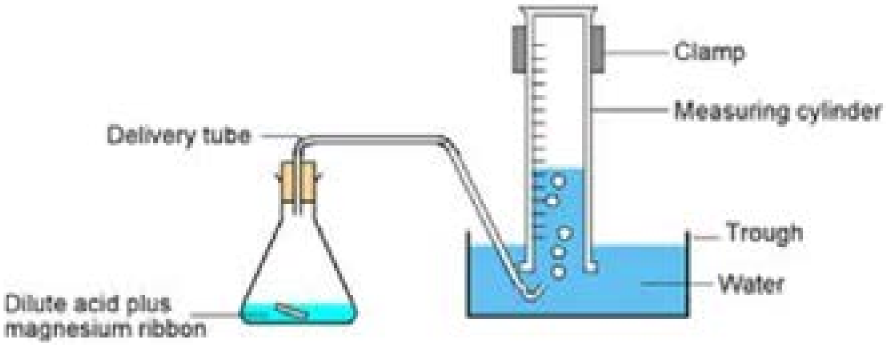
e.g. magnesium + sulfuric acid ⟶ magnesium sulfate + hydrogen

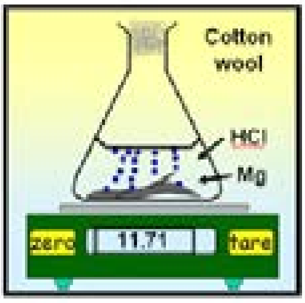
$Mg\ (s) + H_2SO_4\ (aq) \longrightarrow MgSO_4\ (aq) + H_2\ (g)$

e.g. zinc + hydrochloric acid ⟶ zinc chloride + hydrogen

$Zn\ (s) + 2HCl\ (aq) \longrightarrow ZnCl_2\ (aq) + H_2\ (g)$

This experiment, with any metal and acid, can be carried out in two ways to obtain results which can be used to calculate the rate of the reaction. The volume of hydrogen gas given off over time can be measured using the apparatus shown below or the loss of hydrogen gas from the flask on a balance can be measured over time.





Method for balance reaction:

- Measure a volume of dilute acid into the conical flask.
- Have a loose plug of cotton wool in the neck of the flask to prevent "spitting" of liquid droplets which would reduce the accuracy of the results obtained and potentially cause injury.
- Use a balance to weigh a short length of magnesium ribbon.
- Add the magnesium to the flask, replace the cotton wool in the neck and start to record mass readings from the balance at 30 second intervals.

2. Decomposition of hydrogen peroxide using manganese dioxide as a catalyst

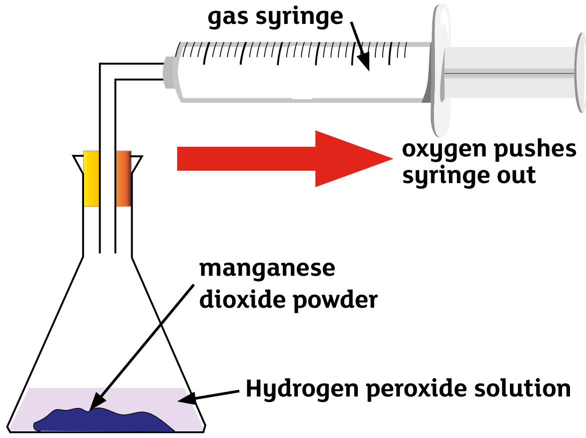
Hydrogen peroxide is produced in our bodies and would poison us if it was not dealt with quickly. In the liver there are enzymes (biological catalysts) which break it down into harmless products.

In the laboratory, hydrogen peroxide decomposes very slowly to form water and oxygen gas:

hydrogen peroxide ⟶ water + oxygen

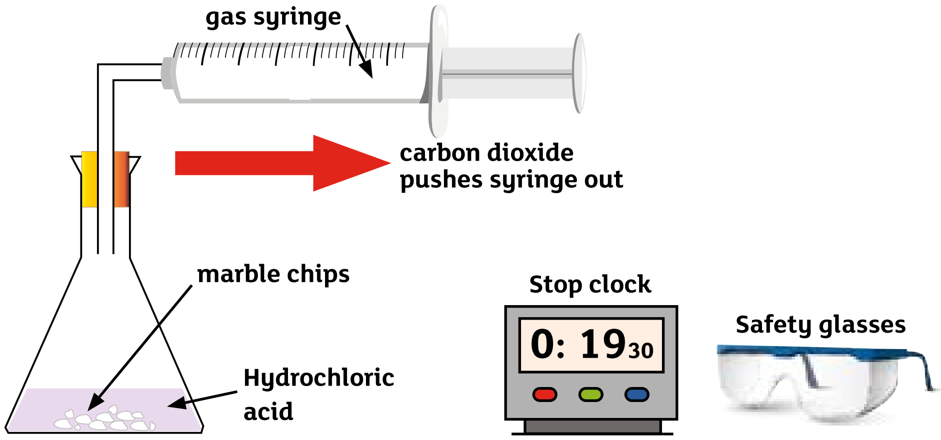
$2H_2O_2 \longrightarrow 2H_2O + O_2$

This reaction can be speeded up by adding a black powder called manganese(IV) oxide (also called manganese dioxide) which acts as a catalyst for this reaction. Lots of fizzing can be observed as oxygen gas is quickly given off.



3. Calcium carbonate and hydrochloric acid reaction

Calcium carbonate (often in the form of marble chips) reacts with acid to give off carbon dioxide gas, which can be collected in a gas syringe. Volume readings of the gas produced can be taken at 30 second intervals using the same apparatus shown in reaction 1.



metal carbonate + acid ⟶ metal salt + water + carbon dioxide

calcium carbonate + hydrochloric acid ⟶ calcium chloride + water + carbon dioxide

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$$

Method:

- Measure a volume of dilute acid into the conical flask.
- Set up the syringe, flask and connector.
- Use a balance to weigh an agreed mass of marble chips.
- Add the marble chips to the flask, replace the rubber bung in the neck and start to record volume readings from the syringe at 30 second intervals.

4. Sodium thiosulfate and acid (equation not required);

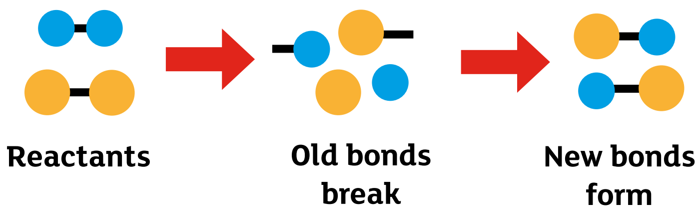
When sodium thiosulphate reacts with an acid, a **yellow precipitate** of sulfur is formed which makes the colourless solution turn cloudy.

hydrochloric acid + sodium thiosulfate ⟶ sodium chloride + sulfur dioxide + sulfur + water

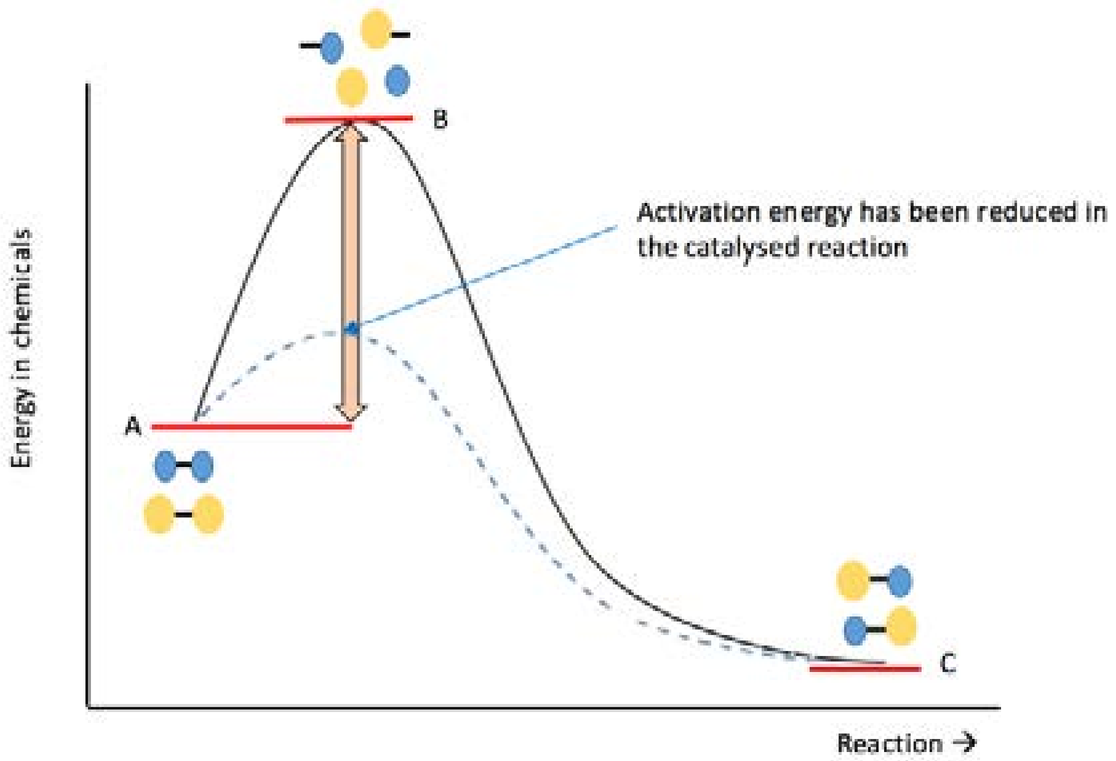
Sulphur dioxide is produced during the experiment, which is very soluble in water, therefore no bubbles of gas are given off. The sodium chloride salt is also soluble therefore the only way to mark the progress of this reaction is to measure the rate of formation of insoluble sulfur. This can be carried out by marking a black x on a piece of white paper before allowing the reaction to take place in a conical flask, placed over the x. As the reaction progresses, and the production of sulfur increases, the ability to see the x through the cloudy precipitate decreases. When carrying out the experiment, use a stop watch to record how long it takes for the x to disappear whilst changing, for example, the concentration of acid used or the temperature at which the reaction is carried out.

How much energy do particles need in order to be able to react?

Most reactions are exothermic (give out heat) overall but there is still a need for energy input to get the reaction started.



This minimum amount of energy that reacting particles must possess is called the **activation energy**. This energy is used to break bonds in the reacting substances so that new products can form.



The bonds in the reacting particles need to be broken before a reaction can take place. A certain level of energy input is required in order to do this, therefore bond breaking is an endothermic process. The amount of energy required to break the bonds is represented on the graph by the arrow between A and B. This is known as the activation energy. Once the bonds in the reactants have been broken, the new bonds in the products can then be formed, shown at point C on the graph. Adding a catalyst to a reaction provides a reaction pathway with lower activation energy. This is represented by the dotted line in the graph above.

A catalyst can be defined as a substance which provides an alternative reaction pathway of lower activation energy or as a substance which increases the rate of a chemical reaction without being used up / remains chemically unchanged at the end.

Revision Questions

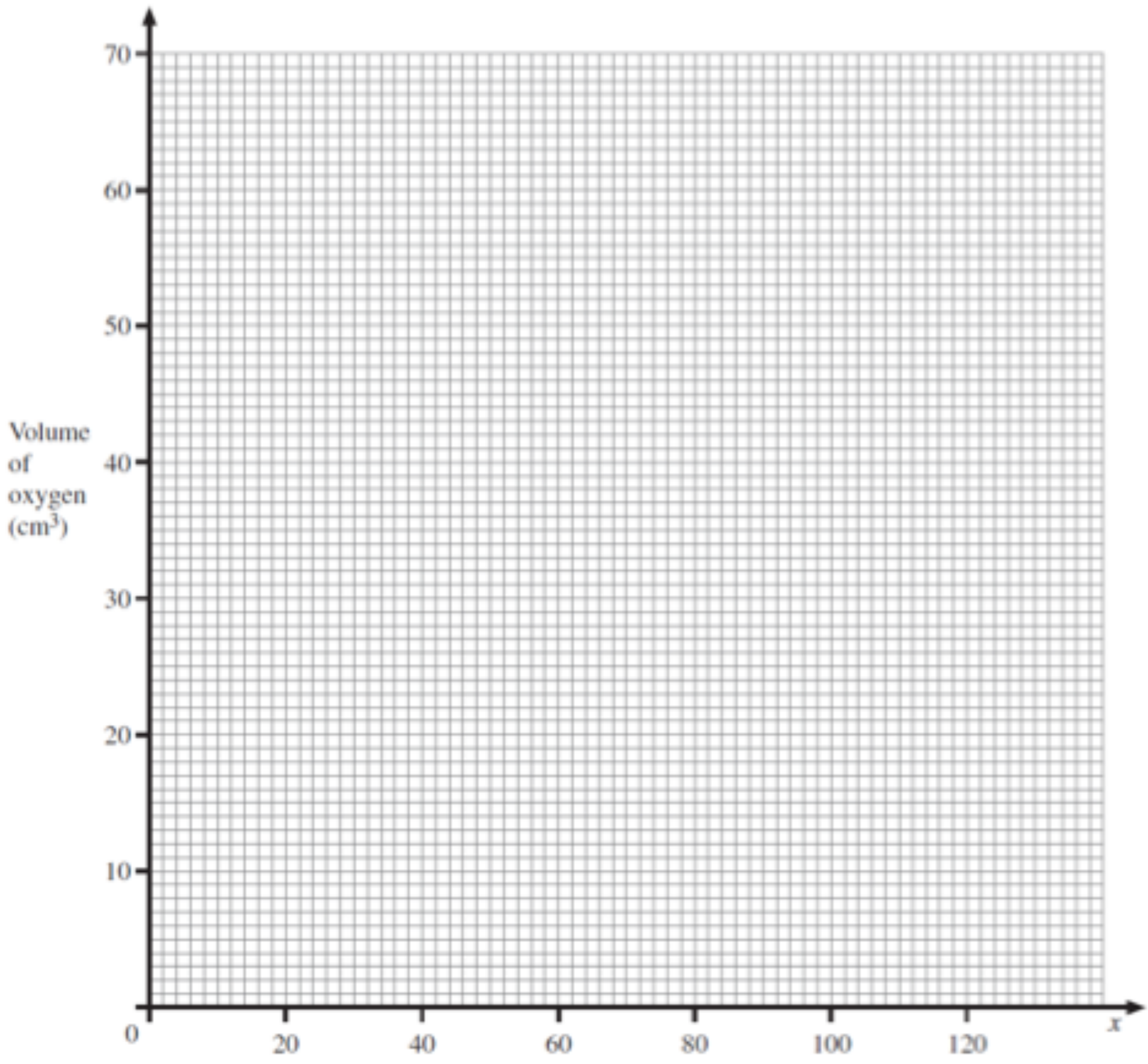
1. (a) The rate of a chemical reaction can be changed when the conditions are altered. Complete the table below to show whether the reaction rate increases, decreases or remains the same. One has been done for you.

New condition	**Effect on rate** Complete the table by writing increase, decrease or none in each box
higher pressure	increase
higher concentration	
lower temperature	
solid particles made smaller	
catalyst added	

[4]

(b) Oxygen gas can be prepared in the laboratory by the decomposition of hydrogen peroxide using manganese(IV) oxide as the catalyst. The results given in the table below were obtained when 1 g of manganese(IV) oxide was added to 50 cm^3 of hydrogen peroxide at 25 °C.

Time (s)	0	20	40	60	80	100	120
Volume of oxygen (cm^3)	0	30	49	59	63	64	64



(i) Label the x-axis above. [1]

(ii) On the grid above plot a curve to show the results. [3]

(iii) From the graph, how long will it take for 56 cm^3 of oxygen to be formed?

____ [1]

(iv) Use the idea of collision to explain the effect of increasing the temperature on the rate of this reaction.

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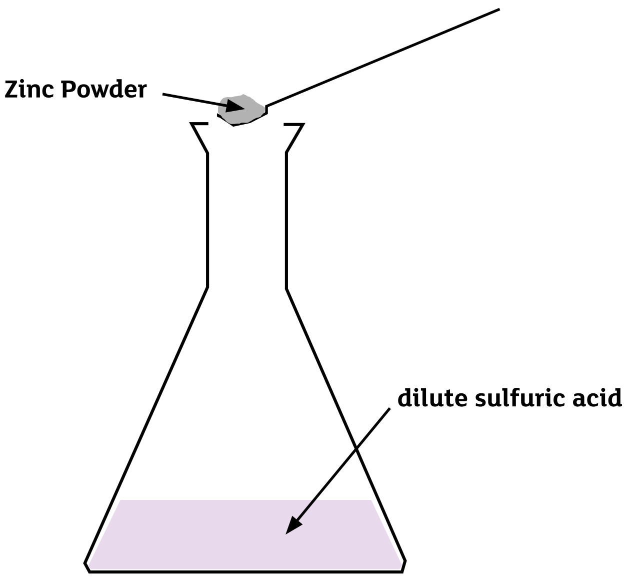
____ [3]

(c) A student wanted to investigate the rate of the reaction between zinc and sulfuric acid. He put 1 g zinc powder into a conical flask containing 50 cm^3 dilute sulfuric acid and shook the flask carefully. After 10 seconds the reaction had finished.



(i) Give **three** ways in which the student could change the investigation, using the same mass of zinc, to **slow down** the reaction.

1. ____

2. ____

3. ____ [3]

(ii) How could the student tell when the reaction of zinc with dilute sulfuric acid had ended?

____ [1]

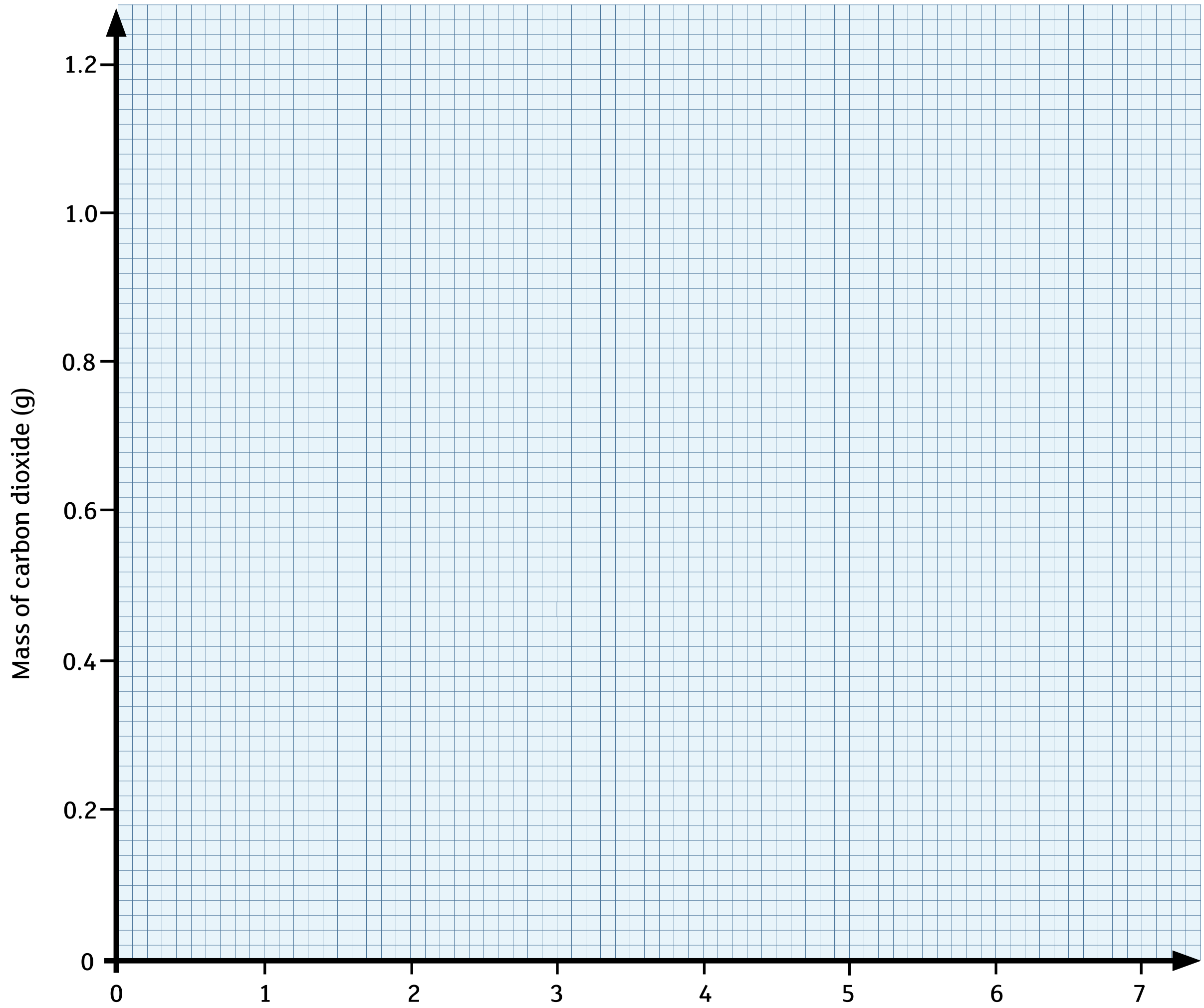
2. (a) Marble chips react with dilute hydrochloric acid as shown below:

$$CaCO_{3\,(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2\,(aq)} + H_2O_{(l)} + CO_{2\,(g)}$$

A student measured the mass of carbon dioxide being produced over a period of time when very small marble chips were added to dilute hydrochloric acid. The marble chips were in excess.

The results are shown in the table below.

Mass of carbon dioxide (g)	0.0	0.40	0.68	0.85	0.96	1.02	1.04	1.04
Time (mins)	0	1	2	3	4	5	6	7



(i) Label the x-axis on the grid above. [1]

(ii) On the grid above plot a curve to show the results. [3]

(iii) At what time did the reaction stop?

____ [1]

(iv) From your graph, how long did it take for 0.5 grams of carbon dioxide to be formed?

____ [1]

(v) Use the idea of collisions to explain the effect of **increasing** the size of the marble chips on the rate of reaction.

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____ [3]

(vi) The student repeated the reaction using the same volume and concentration of dilute hydrochloric acid as before but this time using larger marble chips. How would using larger marble chips affect the **total mass** of carbon dioxide formed?

____

____ [1]