Lesson Six: Rates of Reaction

Aims

By the end of this lesson you should be able to:

- describe the effects on rates of reaction of:
    - surface area of a solid
    - concentration of a solution
    - pressure of a gas
    - temperature
    - use of a catalyst
- explain the effects on rates of reaction, in terms of particle collision theory, of:
    - surface area of a solid
    - concentration of a solution
    - pressure of a gas
    - temperature
- describe experiments to investigate the effects on rates of reaction of:
    - surface area of a solid
    - concentration of a solution

  (both to include the reaction between marble chips and dilute hydrochloric acid)
    - temperature
    - use of a catalyst (to include the catalytic decomposition of hydrogen peroxide solution)
- know that a catalyst:
    - increases the rate of a reaction, but is chemically unchanged at the end of the reaction
    - works by providing an alternative pathway involving a lower activation energy
- **draw and explain reaction profile diagrams showing ΔH and activation energy**

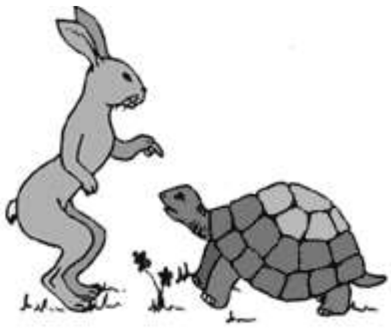
Context

This lesson covers sections 3.9 – 3.16 of the Edexcel IGCSE Chemistry specification.

Edexcel International GCSE (9-1) Chemistry Student Book, pages 227-237.

Introduction

The **rate** of a chemical reaction means the speed at which it goes. Some chemical reactions take place very slowly: for instance, a bone may take millions of years to turn into a fossil. Other reactions take place within a fraction of a second: for instance an explosion. The longer the time a reaction takes to finish, the slower the rate of the reaction.



A chemical reaction involves reactants turning into products:

reactants ----> products

As the reaction progresses, the reactants are used up while the products are made.

To measure the *rate* of a reaction we can measure one of two things:

- how fast a reactant is used up, or
- how fast a product is made

Depending upon the substance measured, we might express a rate of reactions in a variety of different units, such as g/s (grams per second) or cm^3/min (centimetres cubed per minute). The first half of the unit will always be a measure of the “amount” used up or made, while the second half will be a measure of time in which this happens.

In this lesson we look at four different things:

- factors that affect rates of reaction
- reasons that these factors have their effect
- how rates can be measured experimentally
- how energy changes during a reaction can be shown using a diagram.

This information will then be used in Lessons Seven to Nine to plan, carry out and interpret an investigation into the rate of one particular reaction.

Particle Collision Theory

As we learned in Lesson One, all substances are made of **particles** that move in various ways.

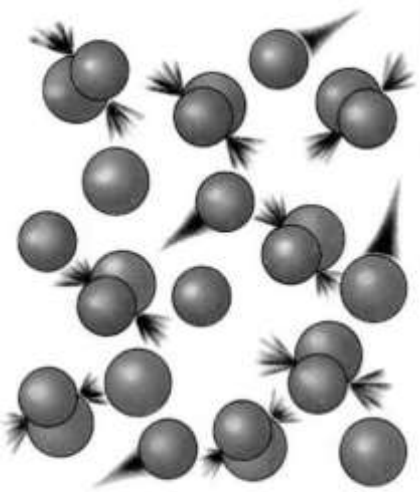
Consider the chemical reaction:

A + B ----> C + D

The particle collision theory says that A and B will only react to give C and D:

- when a particle of A and a particle of B **collide**
- and if the colliding particles have sufficient energy, called the **activation energy** of the reaction

This theory can be used to explain the effect of various factors upon rates of reactions.



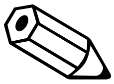
Log on to Twig and look at the film titled: **Rates of Reaction: Basics**

www.ool.co.uk/1451by

Knowing the rate of chemical reactions can be very important, particularly in industry. Discover which factors indicate the rate of a chemical reaction, and how you can change it.

Factors affecting Rates of Reactions

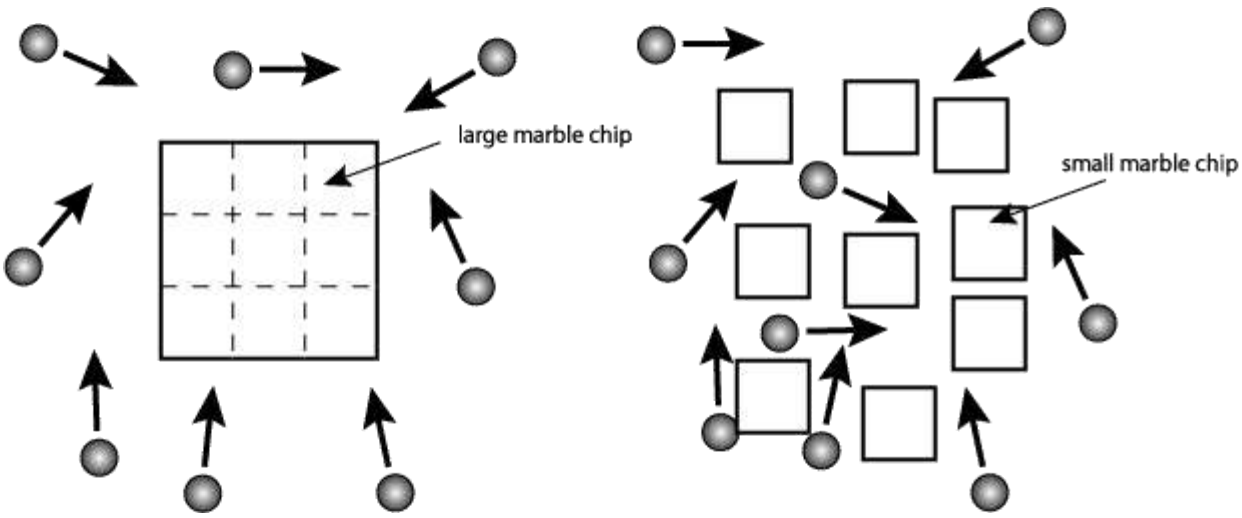
1. Surface Area of a Solid

Activity 1	Investigating the effect of the surface area of a solid
	**Materials and equipment**: — 2 glasses, 2 soluble aspirin tablets, cold water. **Method** — Take 2 glasses of cold water. Crush up one of the aspirin tablets, but leave the other as a whole tablet. Add one to each glass of water at the same time and watch. **Result** — The crushed aspirin dissolves much faster.

If one reactant in a reaction is a solid, and the other is a liquid or gas, increasing the surface area of the solid makes the reaction go faster.

It is easy to see why this is so. Consider acid (a liquid) reacting with a block of marble (a solid). The reaction can only take place at the surface of the block, because this is the only place where marble particles and acid particles can *collide*.

If you cut or grind the block up into smaller pieces, the marble surface exposed to the acid increases, as shown below. This means more marble-particle acid-particle collisions can take place per second, so the reaction speeds up.



The small circles represent the acid particles moving around in the liquid

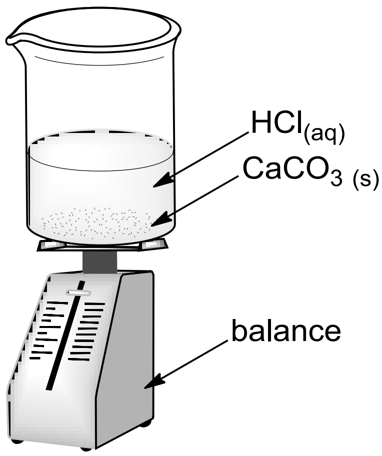
This can be investigated in a laboratory using marble chips (chemical name **calcium carbonate**) and **dilute hydrochloric acid** (hydrochloric acid mixed with a lot of water). See also Activity 5 and Activity 6 on pages 228-231 of the textbook.

The word and chemical equations for this reaction are as follows:

Calcium carbonate + hydrochloric acid ➔ calcium chloride + water + carbon dioxide

$$CaCO_{3}\,_{(s)} + 2HCl_{(aq)} \dashrightarrow CaCl_{2}\,_{(aq)} + H_{2}O_{(l)} + CO_{2}\,_{(g)}$$

As carbon dioxide is a gas, it forms bubbles (bubbling is called **effervescence**) in the liquid as it is formed, and these escape. As this results in a loss of **mass** from the reaction mixture, we can follow the reaction by weighing the beaker containing the substances on a balance. The reading on the **balance** goes down as the carbon dioxide escapes, and the speed with which it goes down is a measure of the rate of the reaction. An **excess** of acid is used, to make sure that all of the marble will react.



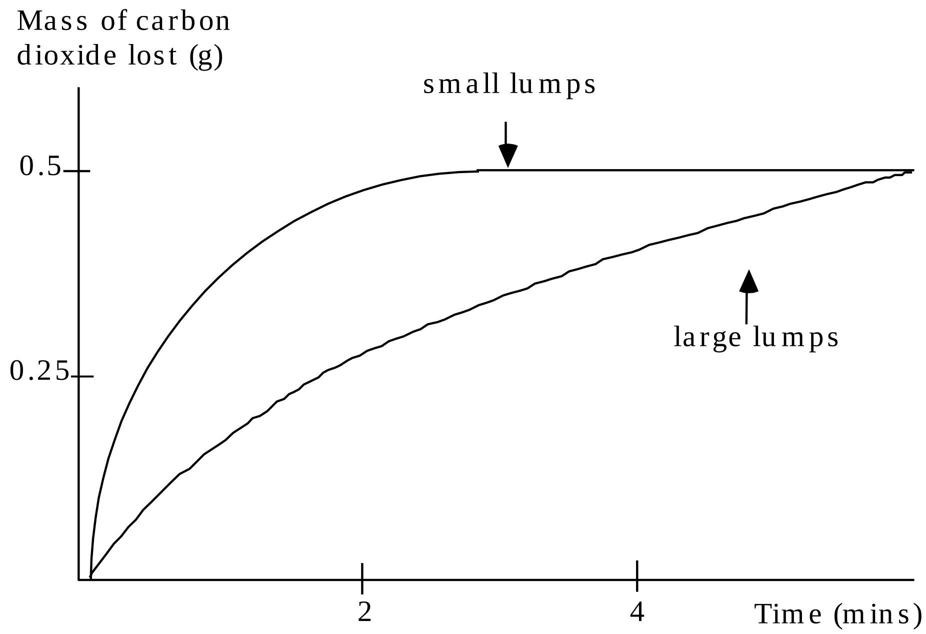
The results from this investigation can be presented in a table like this:

Mass reading on balance (g)	Mass of carbon dioxide given off (g)	Time (min)
98.56	0.00	0.0
97.31	1.25	0.5
96.59	1.97	1.0

Notice that the mass of carbon dioxide produced is a **cumulative total** (it adds up as you continue).

The same experiment is then repeated under identical conditions, except that the marble is crushed into small pieces first. This makes the reaction go faster.

The results can be presented in a graph like the one below:



Notice the following points:

- The final amount of carbon dioxide formed in both cases is the same, because the mass of calcium carbonate used is the same. All of the calcium carbonate reacts eventually in both cases because the acid is **in excess** (there is more than enough of it to react with all of the calcium carbonate).

- The graph is steeper at the beginning if the lumps are smaller. The slope of the graph tells us the **rate** of the reaction: the steeper the **gradient** of the line, the greater the rate of reaction

- As time goes on, the acid particles are used up, so the concentration of the acid solution drops. That is why the gradient of the curve decreases with time. The time at which the concentration of the acid is the same in the two experiments is time zero, so it is the gradient of the curves at time zero that should be compared.

This experiment confirms that increasing the surface area of a solid reactant speeds up the rate of the reaction.

2. Concentration of a Solution

A **solution** is formed by dissolving a solid or gas (the **solute**) in a liquid (the **solvent**). See Lesson One. Usually at IGCSE the liquid involved is water, and the solution is called an **aqueous solution** (***aq***). See Lesson Five.

A solution is a mixture of solute and solvent particles moving around at random. The **concentration** of a solution is a measure of the mass of solute dissolved in a given amount of solution. As the concentration of a solution goes up, there are more solute particles per cm^3 of the solution: they become more crowded together.

Now consider reactions where:

- a solution reacts with a solid, or

- two solutions react together.

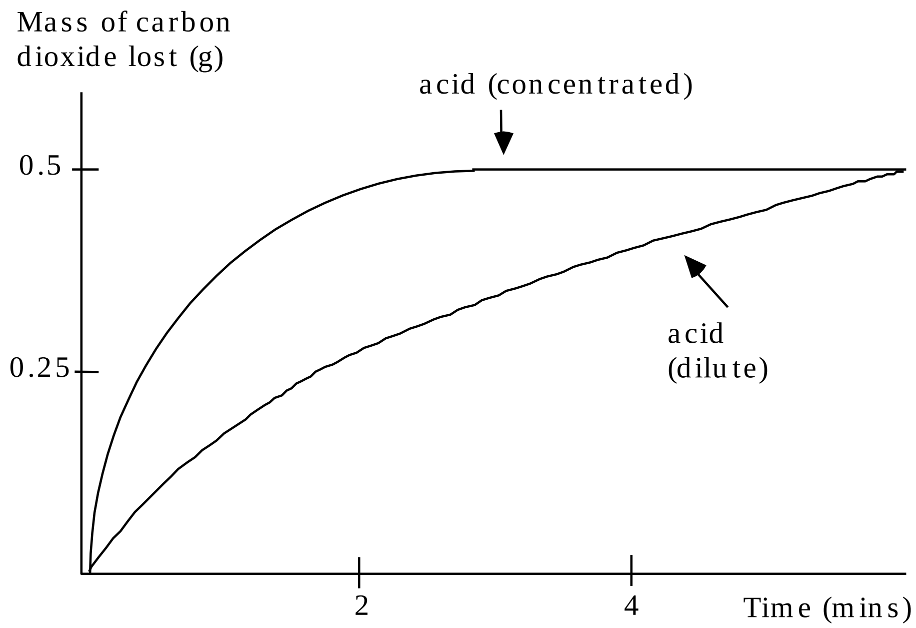
In both cases, if a solution becomes more concentrated, there will be more collisions per second between reactant particles, and the rate of reaction increases.

We will look at two examples of this, one of a solution reacting with a solid, and the other of two solutions reacting together.

(a) Marble chips and hydrochloric acid

This is the same reaction that we looked at in the previous section. See also Activity 7 on page 233 of the textbook.

Two experiments, each using 2g of large marble chips and excess acid, are carried out. In the first, 50cm^3 of dilute hydrochloric acid is used. In the second, 50cm^3 of more concentrated hydrochloric acid is used. The results are shown in the graph below:



As before, the gradient (slope) of the graph gives the rate of each reaction. As expected, at the beginning the marble chips react faster with the concentrated acid.

Activity 2	Put "rates of reaction concentration" into the YouTube search box at www.youtube.com to see some demonstrations of this experiment.

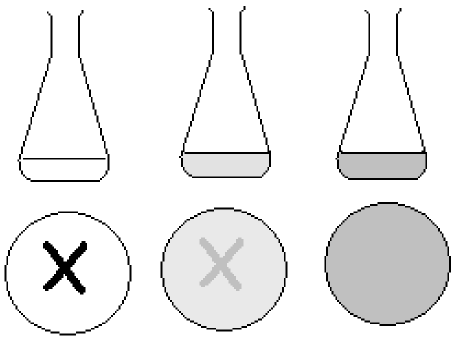
(b) Sodium thiosulphate and hydrochloric acid

When an aqueous solution of sodium thiosulphate is mixed with dilute hydrochloric acid, the following reaction takes place:

$Na_2S_2O_3(aq) + 2HCl(aq) \dashrightarrow 2NaCl(aq) + S(s) + SO_2(g) + H_2O(l)$

The solid sulphur (S) forms as a pale yellow powder **suspended** in the liquid, making it go cloudy. The more sulphur formed, the more cloudy the liquid becomes.

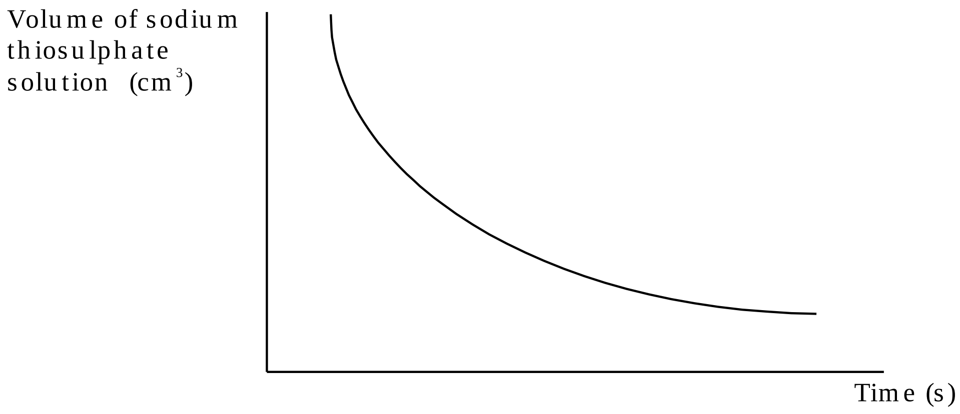
Dilute hydrochloric acid is mixed with sodium thiosulphate solution in a glass **conical flask**, as shown below, and the flask is placed over a piece of paper on which a cross has been drawn. We measure the time taken before the cloudiness stops the cross being seen when viewed from above through the liquid. The shorter the time, the faster the rate of reaction, and vice versa.



The reaction is carried with different concentrations of sodium thiosulphate solution, while keeping everything else (e.g. temperature) the same. The results are shown below:

	Volume of sodium thiosulphate solution (cm^3)	Volume of water (cm^3)	Time taken to obliterate cross (s)
1	50	0	Short time
2	40	10	↓
3	30	20	
4	20	30	
5	10	40	Long time

A graph of the results looks like this:

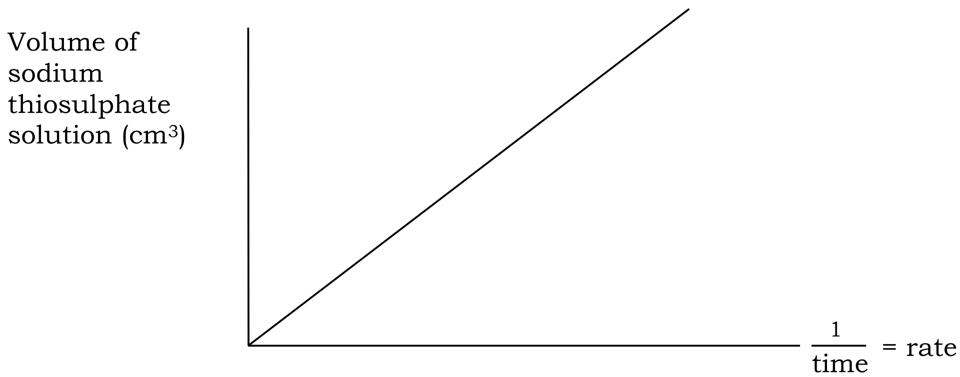


As expected, the greater the volume of sodium thiosulphate solution used to form the mixture (and therefore the greater its *concentration*) the shorter the time for the reaction to finish (so the faster the rate of reaction):

higher concentration ---> faster rate of reaction

Rate of reaction and time to finish

Because rate of reaction increases as time taken to finish decreases, the *rate of reaction can be represented as 1/time* (one divided by time). If a graph of this is plotted against concentration, it becomes a straight line passing through the origin of the graph like this:



At the origin, the concentration of the solution is zero, so the time for the reaction to finish is infinite. One divided by infinity is zero, so in this case 1/time is also zero.

Activity 3

Put "sodium thiosulphate hydrochloric acid" into the YouTube search box at www.youtube.com to see some demonstrations of this experiment, investigating the effect of either temperature or concentration on the rate of the reaction.

3. Temperature

Activity 4

Take two glasses, one of cold water and one of warm water. Place a sugar cube into each. Do not stir the water in either of the glasses, but note how long it takes each to dissolve. The sugar should dissolve faster in the warm water.

When the temperature of a substance is increased, the average speed of movement of its particles increases, which means that on average each of them has more energy. This makes the rate of a chemical reaction increase with increasing temperature for two reasons:

- the particles of reactants collide more frequently, as they are moving faster
- when the particles do collide, they are more likely to have sufficient energy to react

This can be investigated using either of the two reactions described above. The temperatures of the liquids are measured using a **thermometer**.

4. Pressure of Gases

If you increase the **pressure** of a gas, you push its particles closer together. This increases the rate of reaction, if one or more of the reactants is a gas, for the same reasons that increasing the concentration of a solution does: there will be more collisions of reactant particles per second.

5. Use of a Catalyst

A **catalyst** is a substance which *increases the rate of a reaction but is chemically unchanged at the end of the reaction.* It does not appear as a reactant or product in the equation: instead it is written over the top of the equation's arrow.

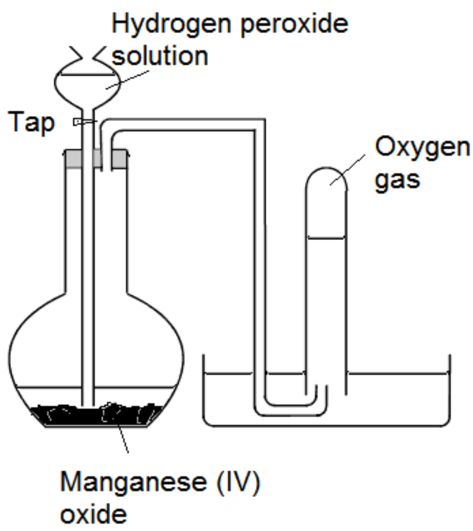
An example, often used in school laboratories, is the speeding up of the decomposition (breakdown) of hydrogen peroxide into oxygen and water by the catalyst manganese (IV) oxide (MnO_2):

Hydrogen peroxide ➔ water + oxygen

$$2H_2O_{2\,(aq)} \longrightarrow 2H_2O_{(l)} + O_{2\,(g)}$$

Hydrogen peroxide is used as a colourless dilute solution in water, while manganese (IV) oxide is a black power. The oxygen forms bubbles in the liquid as it is given off.

Using the following apparatus, the volume of oxygen gas produced each minute can be measured:



Small weighed amounts of the catalyst are added to the conical flask, and the volume of oxygen released in 1 minute is measured. The volume of the gas produced is **directly proportional** to the mass of the catalyst added. That is, if you double the mass of the catalyst, you double the volume of gas produced per minute.

If you remove the catalyst at the end, dry it and reweigh it, you find that its mass is unchanged. The catalyst is not used up or made during the reaction.

N.B. The other factors which can change the rate of this reaction must be kept constant to make it a fair test. These include:

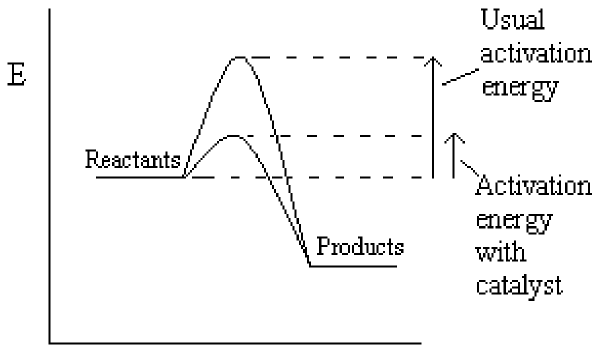
- concentration and volume of the hydrogen peroxide solution
- temperature
- size of each small lump of manganese (IV) oxide powder.

A more accurate way of measuring the volume of the oxygen produced, using a **gas syringe**. See Activity 8 on page 236 of the textbook.

Activity 5	Put "rates of reaction catalyst" into the YouTube search box at www.youtube.com to see some demonstrations of this experiment.

Activation Energy

The diagram below is called a **reaction profile**. It shows energy changes as a reaction proceeds, going from left to right.



In this particular reaction, the products have less energy than the reactants (they are lower down on the diagram), so energy is released. Such reactions are said to be **exothermic**. Burning a match is like this – energy is released in the form of heat.

The amount of energy released in given the symbol **ΔH**. It is a negative number for exothermic reactions. We will investigate this more in Lesson 16.

Even though energy will be released as a result of the reaction, you still have to give the reactants some extra energy to get the reaction started. This is represented by the “hump” on the diagram. This extra energy is called the **activation energy** for the reaction. Often this is supplied in the form of heat, which is what happens when you strike a match on a matchbox to light it.

Catalysts and Activation Energy

Remember: to react, two reactant particles must collide with sufficient energy. The minimum necessary amount of energy is the activation energy.

In any mixture, the particles will have a range of energies, because they will be moving at a range of speeds. *Only collisions between the most energetic will have at least the activation energy and will be fruitful.* Other, lower energy, collisions will not be fruitful: the reactant particles will just bounce off each other unchanged.

A catalyst works by providing an alternative pathway (route) for the reaction which has *a lower activation energy*. This means that a higher percentage of the collisions will have enough energy and will be fruitful. So the catalyst speeds the reaction up.

Because catalysts speed up up reactions without having to raise the temperature, and because raising the temperature costs money in the form of fuel, catalysts are very important in industry.

Activity 6

Explain what is wrong with these statements:

(a) Increasing the concentration of a solution increases the rate of reaction because the particles move faster.

(b) Raising the temperature makes a reaction go faster because it reduces the activation energy.

(c) The rate of a reaction is directly proportional to the time it takes to finish.

(d) Limestone bubbles and gives off carbon dioxide when the catalyst hydrochloric acid is added to it.

Now read pages 227-237 of your textbook to consolidate your knowledge and understanding of this lesson.

Keywords

rate of reaction	**concentration**
activation energy	**effervescence**
balance	**surface area**
pressure	**catalyst**
concentration	**gradient**
directly proportional	**solute**
reaction profile	**solvent**
conical flask	**solution**
thermometer	**suspended/suspension**
exothermic	**gas syringe**

Summary

Lesson Six

Particle collision theory

Factors affecting rate

- surface area of solid
- concentration of solution
- pressure of gas
- temperature
- catalysts

Activation energy

- reaction profile
- catalysts

What you need to know

- The meanings of the terms in **bold** in this lesson
- How surface area, concentration, temperature, pressure and catalysts affect rates of reaction
- What catalysts are and what they do

What you might be asked to do

- Explain the effects of surface area, concentration, temperature and pressure on rates of reaction in terms of particle collision theory
- Explain how catalysts speed up reactions
- Interpret the results of experiments (including tables and graphs) concerning rates of reaction

Suggested Answers to Activities

Activity 6

(a) The particles move at the same speed, but they are closer together so there are more collisions which speeds up the reaction.

(b) The activation energy remains the same. However more of the particles have at least the activation energy, so more of the collisions are fruitful, which increases the rate of reaction.

(c) The rate of reaction is directly proportional to *1*/the time it takes to finish (one divided by ...)

(d) The hydrochloric acid is not a catalyst, because it is used up in the reaction. It is a reactant.

Tutor-Marked Assessment B

Write the answers to these questions on file paper, or word-process and print them, and send them to your Tutor for marking. Do not send the questions as well.

You are advised to look back carefully over Lessons 4-6, and the associated reading in the textbook, to research the answers to the questions. Do not just attempt them from memory.

Question 1

Give the correct technical term which matches each of the following definitions:

(a) Able to bend without breaking, like a metal

(b) The substances formed in a chemical reaction

(c) The mass of a solid contained in a certain amount of a solution

(d) The giant three-dimensional structure formed by ions in an ionic substance

(e) The bubbling seen when a gas is formed in a liquid

(f) The attractive force that holds two molecules together in a simple covalent substance

(g) The minimum energy needed by colliding particles if they are going to react

(h) An electron which is not tightly bound to a particular atom, but is able to move throughout a material

(i) The solid which is dissolved in a solution

(j) The shorthand which says if a substance is solid, liquid, gas, or dissolved in water.

(10 marks)

Question 2

Explain the following fully:

(a) Diamond and candle wax are both substances whose atoms are held together by covalent bonds. Yet diamond is much harder, and melts at a much higher temperature, than candle wax. (6)

(b) A chemical reaction goes faster if an appropriate catalyst is present. (4)

(Total: 10 marks)

Question 3

Use the information in Appendices A and B at the back of your file to help you answer the following:

(a) Work out the chemical formulae for the following substances:

(i) potassium chloride (1)

(ii) magnesium carbonate (1)

(iii) sodium sulphate (1)

(iv) iron(III) oxide (1)

(v) silicon oxide (1)

(vi) nitrogen hydride (ammonia) (1)

(b) Suggest the names of the following compounds:

(i) Al_2S_3 (1)

(ii) $MgSO_4$ (1)

(iii) $Ba(NO_3)_2$ (1)

(c) Balance the following chemical equations:

(i) $Ca + O_2 \text{--->} CaO$ (2)

(ii) $C_2H_6 + O_2 \text{--->} CO_2 + H_2O$ (2)

(iii) $Fe_2O_3 + CO \text{--->} Fe + CO_2$ (2)

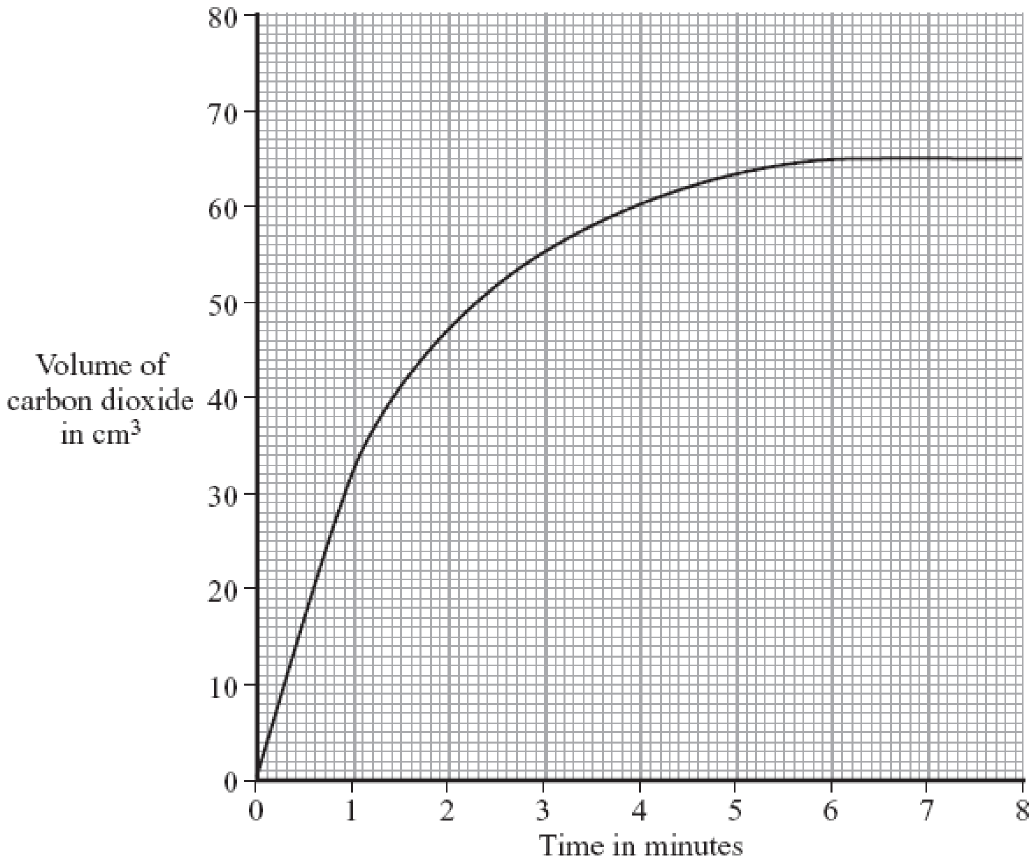
(Total: 15 marks)

Question 4

A student reacted dilute hydrochloric acid with an excess of powdered calcium carbonate at 20^0C. Here is the word equation for the reaction:

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

She measured the volume of carbon dioxide produced in the experiment, and her results are shown in the graph below:



(a) After how many minutes had all the acid been used up? (1 mark)

(b) (i) What happened to the rate of the reaction between 1 minute and 4 minutes? (1 mark)

(ii) Explain the reason for this change. (3 marks)

(c) The experiment was repeated at 35^0C, but everything else was kept the same. Compare the graph that would be obtained with the original graph. (2 marks)

(d) In a third experiment, the same mass of calcium carbonate lumps was used instead of powder at 20^{0}C.

(i) What would this do to the rate of the reaction in the first minute? (1 mark)

(ii) What would this do to the total volume of carbon dioxide released during the reaction? (1 mark)

(Total: 9 marks)

Question 5

The table shows the properties of four substances.

Use the information in the table to answer the following questions.

Substance	Melting point in °C	Boiling point in °C	Conducts electricity when	
			solid	liquid
A	1650	2230	no	no
B	1538	2862	yes	yes
C	−7	59	no	no
D	801	1413	no	yes

Choose from A-D the substance that is most likely to be:

(a) a metal (1)

(b) a giant covalent structure (1)

(c) an ionic compound (1)

(d) a liquid at 25^{0}C (1)

(e) a liquid at 1600^{0}C (1)

(f) a simple covalent substance (1)

(Total: 6 marks)

Total marks for TMA = 50