10.1 Rates of reaction

Fast and slow
Some reactions are fast and some are slow. Look at these examples:

- The precipitation of silver chloride, when you mix solutions of silver nitrate and sodium chloride. This is a very fast reaction.
- Concrete setting. This reaction is quite slow. It will take a couple of days for the concrete to fully harden.
- Rust forming on an old car. This is usually a very slow reaction. It will take years for the car to rust completely away.

But it is not always enough to know just that a reaction is fast or slow. In factories where they make products from chemicals, they need to know exactly how fast a reaction is going, and how long it will take to complete. In other words, they need to know the rate of the reaction.

What is rate?
Rate is a measure of how fast or slow something is. Here are some examples.

- This plane has just flown 800 kilometres in 1 hour. It flew at a rate of 800 km per hour.
- This petrol pump can pump out petrol at a rate of 50 litres per minute.
- This machine can print newspapers at a rate of 10 copies per second.

From these examples you can see that:
Rate is a measure of the change that happens in a single unit of time.
Any suitable unit of time can be used – a second, a minute, an hour, even a day.
When zinc is added to dilute sulfuric acid, they react together. The zinc disappears slowly, and a gas bubbles off.

As time goes by, the gas bubbles off more and more slowly. This is a sign that the reaction is slowing down.

Finally, no more bubbles appear. The reaction is over, because all the acid has been used up. Some zinc remains behind.

The gas that bubbles off is hydrogen. The equation for the reaction is:

\[ \text{zinc} + \text{sulfuric acid} \rightarrow \text{zinc sulfate} + \text{hydrogen} \]
\[ \text{Zn (s)} + \text{H}_2\text{SO}_4 \text{(aq)} \rightarrow \text{ZnSO}_4 \text{(aq)} + \text{H}_2 \text{(g)} \]

Both zinc and sulfuric acid get used up in the reaction. At the same time, zinc sulfate and hydrogen form.

You could measure the rate of the reaction, by measuring:
- the amount of zinc used up per minute or
- the amount of sulfuric acid used up per minute or
- the amount of zinc sulfate produced per minute or
- the amount of hydrogen produced per minute.

For this reaction, it is easiest to measure the amount of hydrogen produced per minute, since it is the only gas that forms. It can be collected as it bubbles off, and its volume can be measured.

**In general, to find the rate of a reaction, you should measure:**
- the amount of a reactant used up per unit of time or
- the amount of a product produced per unit of time.

1. Here are some reactions that take place in the home. Put them in order of decreasing rate (the fastest one first).
   a. raw egg changing to hard-boiled egg
   b. fruit going rotten
   c. cooking gas burning
   d. bread baking
   e. a metal tin rusting

2. Which of these rates of travel is slowest?
   - 5 kilometres per second
   - 20 kilometres per minute
   - 60 kilometres per hour

3. Suppose you had to measure the rate at which zinc is used up in the reaction above. Which of these units would be suitable? Explain your choice.
   - a. litres per minute
   - b. grams per minute
   - c. centimetres per minute

4. Iron reacts with sulfuric acid like this:
   \[ \text{Fe (s)} + \text{H}_2\text{SO}_4 \text{(aq)} \rightarrow \text{FeSO}_4 \text{(aq)} + \text{H}_2 \text{(g)} \]
   a. Write a word equation for this reaction.
   b. Write down four different ways in which the rate of the reaction could be measured.
10.2 Measuring the rate of a reaction

A reaction that produces a gas
The rate of a reaction is found by measuring the amount of a reactant used up per unit of time, or the amount of a product produced per unit of time. Look at this reaction:

\[
\text{magnesium} + \text{hydrochloric acid} \rightarrow \text{magnesium chloride} + \text{hydrogen}
\]

\[
\text{Mg} (s) + 2\text{HCl} (aq) \rightarrow \text{MgCl}_2 (aq) + \text{H}_2 (g)
\]

Here hydrogen is the easiest substance to measure, because it is the only gas in the reaction. It bubbles off and can be collected in a gas syringe, where its volume is measured.

The experiment

Clean the magnesium with sandpaper. Put dilute hydrochloric acid in the flask. Drop the magnesium into the flask, and insert the stopper and syringe immediately. Start the clock at the same time.

Hydrogen begins to bubble off. It rises up the flask and into the gas syringe, pushing the plunger out:

At the start, no gas has yet been produced or collected. So the plunger is all the way in.

Now the plunger has been pushed out to the 20 cm³ mark. 20 cm³ of gas have been collected.

The volume of gas in the syringe is noted at intervals – for example every half a minute. How will you know when the reaction is complete?

Typical results

<table>
<thead>
<tr>
<th>Time / minutes</th>
<th>0</th>
<th>(\frac{1}{2})</th>
<th>1</th>
<th>(\frac{1}{2})</th>
<th>2</th>
<th>(\frac{1}{2})</th>
<th>3</th>
<th>(\frac{1}{2})</th>
<th>4</th>
<th>(\frac{1}{2})</th>
<th>5</th>
<th>(\frac{1}{2})</th>
<th>6</th>
<th>(\frac{1}{2})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume of hydrogen / cm³</td>
<td>0</td>
<td>8</td>
<td>14</td>
<td>20</td>
<td>25</td>
<td>29</td>
<td>33</td>
<td>36</td>
<td>38</td>
<td>39</td>
<td>40</td>
<td>40</td>
<td>40</td>
<td>40</td>
</tr>
</tbody>
</table>

This table shows some typical results for the experiment.

You can tell quite a lot from this table. For example, you can see that the reaction lasted about five minutes. But a graph of the results is even more helpful. The graph is shown on the next page.
Notice these things about the results:

1. In the first minute, 14 cm³ of hydrogen are produced. So the rate for the first minute is 14 cm³ of hydrogen per minute. In the second minute, only 11 cm³ are produced. (25 – 14 = 11) So the rate for the second minute is 11 cm³ of hydrogen per minute. The rate for the third minute is 8 cm³ of hydrogen per minute. So the rate decreases as time goes on. **The rate changes all through the reaction. It is greatest at the start, but decreases as the reaction proceeds.**

2. The reaction is fastest in the first minute, and the curve is steepest then. It gets less steep as the reaction gets slower. **The faster the reaction, the steeper the curve.**

3. After 5 minutes, no more hydrogen is produced, so the volume no longer changes. The reaction is over, and the curve goes flat. **When the reaction is over, the curve goes flat.**

4. Altogether, 40 cm³ of hydrogen are produced in 5 minutes. The average rate for the reaction = \( \frac{\text{total volume of hydrogen}}{\text{total time for the reaction}} \) = \( \frac{40 \text{ cm}^3}{5 \text{ minutes}} \) = 8 cm³ of hydrogen per minute.

Note that this method can be used for any reaction where one product is a gas.

**Q**

1. For the experiment in this unit, explain why:
   - a. the magnesium ribbon is cleaned first
   - b. the clock is started the moment the reactants are mixed
   - c. the stopper is replaced immediately

2. From the graph at the top of this page, how can you tell when the reaction is over?

3. Look again at the graph at the top of the page.
   - a. How much hydrogen is produced in the first:
     - i. 2.5 minutes?
     - ii. 4.5 minutes?
   - b. How long did it take to get 20 cm³ of hydrogen?
   - c. What is the rate of the reaction during:
     - i. the fourth minute?
     - ii. the sixth minute?
10.3 Changing the rate of a reaction (part I)

Ways to change the rate of a reaction
There are several ways to speed up or slow down a reaction. For example you could change the concentration of a reactant, or the temperature. The rate will change – but the amount of product you obtain will not.

1 By changing concentration
Here you will see how rate changes with the concentration of a reactant.

The method  Repeat the experiment from page 128 twice (A and B below). Keep everything the same each time except the concentration of the acid. In B it is twice as concentrated as in A.

The results  Here are both sets of results, shown on the same graph.

Notice these things about the results:
1 Curve B is steeper than curve A. So the reaction was faster for B.
2 In B, the reaction lasts for 60 seconds. In A it lasts for 120 seconds.
3 Both reactions produced 60 cm$^3$ of hydrogen. Do you agree?
4 So in B the average rate was 1 cm$^3$ of hydrogen per second. (60 ÷ 60)
   In A it was 0.5 cm$^3$ of hydrogen per second. (60 ÷ 120)
   The average rate in B was twice the average rate in A.
   So in this example, doubling the concentration doubled the rate.

These results show that:
A reaction goes faster when the concentration of a reactant is increased.
This means you can also slow down a reaction, by reducing concentration.
2 By changing temperature
Here you will see how rate changes with the temperature of the reactants.

The method Dilute hydrochloric acid and sodium thiosulfate solution react to give a fine yellow precipitate of sulfur. You can follow the rate of the reaction like this:
1. Mark a cross on a piece of paper.
2. Place a beaker containing sodium thiosulfate solution on top of the paper, so that you can see the cross through it, from above.
3. Quickly add hydrochloric acid, start a clock at the same time, and measure the temperature of the mixture.
4. The cross grows fainter as the precipitate forms. Stop the clock the moment you can no longer see the cross. Note the time.
5. Now repeat steps 1–4 several times, changing only the temperature. You do this by heating the sodium thiosulfate solution to different temperatures, before adding the acid.

The cross grows fainter with time

The results This table shows some typical results:

<table>
<thead>
<tr>
<th>Temperature / °C</th>
<th>20</th>
<th>30</th>
<th>40</th>
<th>50</th>
<th>60</th>
</tr>
</thead>
<tbody>
<tr>
<td>Time for cross to disappear / seconds</td>
<td>200</td>
<td>125</td>
<td>50</td>
<td>33</td>
<td>24</td>
</tr>
</tbody>
</table>

The cross disappears when enough sulfur has formed to hide it. This took 200 seconds at 20 °C, but only 50 seconds at 40 °C.

So the reaction is four times faster at 40 °C than at 20 °C.

A reaction goes faster when the temperature is raised. When the temperature increases by 10 °C, the rate generally doubles.

That is why food cooks much faster in pressure cookers than in ordinary saucepans. (The temperature in a pressure cooker can reach 125 °C.) And if you want to slow a reaction down, of course, you can lower the temperature.

1. Look at the graph on the opposite page.
   a. How much hydrogen was obtained after 2 minutes in: i. experiment A? ii. experiment B?
   b. How can you tell which reaction was faster, from the shape of the curves?

2. Explain why experiments A and B both gave the same amount of hydrogen.

3. Copy and complete: A reaction goes …… when the concentration of a …… is increased. It also goes …… when the …… is raised.

4. Raising the temperature speeds up a reaction. Try to give two (new) examples of how this is used in everyday life.

5. What happens to the rate of a reaction when the temperature is lowered? How do we make use of this?
3 **By changing surface area**

In many reactions, one reactant is a solid. The reaction between hydrochloric acid and calcium carbonate (marble chips) is an example. Carbon dioxide gas is produced:

\[
\text{CaCO}_3 \, (s) + 2\text{HCl} \, (aq) \rightarrow \text{CaCl}_2 \, (aq) + \text{H}_2\text{O} \, (l) + \text{CO}_2 \, (g)
\]

The rate can be measured using the apparatus on the right.

**The method** Place the marble in the flask and add the acid. Quickly plug the flask with cotton wool to stop any liquid splashing out. Then weigh it, starting the clock at the same time. Note the mass at regular intervals until the reaction is complete.

Carbon dioxide is a heavy gas. It escapes through the cotton wool, which means that the flask gets lighter as the reaction proceeds. So by weighing the flask at regular intervals, you can follow the rate of reaction.

The experiment is repeated twice. Everything is kept exactly the same each time, except the surface area of the marble chips.

For experiment 1, large chips are used. Their surface area is the total area of exposed surface.

For experiment 2, the same mass of marble is used – but the chips are small so the surface area is greater.

**The results** The results of the two experiments are plotted here:

![Graph showing the results of experiments 1 and 2.](image)

So what can you conclude about surface area? Did it affect the rate of the reaction?

**How to draw the graph**

First you have to find the loss in mass at different times:

- loss in mass at a given time = mass at start − mass at that time

Then you plot the values for loss in mass against time.
Notice these things about the results:
1. Curve 2 is steeper than curve 1. This shows that the reaction is faster for the small chips.
2. In both experiments, the final loss in mass is 2.0 grams. In other words, 2.0 grams of carbon dioxide are produced each time.
3. For the small chips, the reaction is complete in 4 minutes. For the large chips, it takes 6 minutes.

These results show that:
**The rate of a reaction increases when the surface area of a solid reactant is increased.**

**Explosion!**
As you have seen, you can increase the rate of a reaction by increasing:
- the concentration of a reactant
- the temperature
- the surface area of a solid reactant

In some situations, an increase in any of these can lead to a dangerously fast reaction. You get an explosion. Here are examples.

**In flour mills**  Flour particles are tiny, so flour has a very large surface area. It can also catch fire. In a flour mill, if there is a lot of flour dust in the air; a spark from a machine could be enough to cause an explosion.

For the same reason, explosions are a risk in wood mills, from wood dust, and in silos where wheat and other grains are stored. And in factories that make custard powder, and dried milk. The dust from all these will burn.

**In coal mines**  In coal mines, methane (CH₄) and other flammable gases collect in the air. At certain concentrations they form an explosive mix with the air. A spark is enough to set off an explosion.

---

**Q**

1. This question is about the graph on the opposite page. For each experiment find:
   - the mass of carbon dioxide produced in the first minute
   - the average rate of production of the gas, for the complete reaction.

2. a. Which has the largest surface area: 1 g of large marble chips, or 1 g of small marble chips?
   b. Which 1 g sample will disappear first when reacted with excess hydrochloric acid? Why?

3. Explain why fine flour dust in the air is a hazard, in flour mills.
10.5 Explaining rates

The collision theory
Magnesium and dilute hydrochloric acid react together like this:

\[
\text{Mg (s) + 2HCl (aq) \rightarrow MgCl}_2 (aq) + \text{H}_2 (g)
\]

In order for the magnesium and acid particles to react together:

- the particles must collide with each other, and
- the collision must have enough energy to be successful. In other words, enough energy to break bonds to allow reaction to occur.

This is called the collision theory. It is shown by the drawings below.

The particles in the liquid move non-stop. To react, an acid particle must collide with a magnesium atom, and bonds must break.

If there are lots of successful collisions in a given minute, then a lot of hydrogen is produced in that minute. In other words, the rate of reaction is high. If there are not many, the rate of reaction is low.

The rate of a reaction depends on how many successful collisions there are in a given unit of time.

Changing the rate of a reaction

Why rate increases with concentration  If the concentration of the acid is increased, the reaction goes faster. It is easy to see why:

In dilute acid, there are not so many acid particles. So there is less chance of an acid particle hitting a magnesium atom.

Here the acid is more concentrated – there are more acid particles. So there is now more chance of a successful collision.

The more successful collisions there are, the faster the reaction.
That idea also explains why the reaction between magnesium and hydrochloric acid slows down over time:

At the start, there are plenty of magnesium atoms and acid particles. But they get used up in successful collisions.

**Why rate increases with temperature** On heating, *all* the particles take in heat energy.

This makes the acid particles move faster – so they collide more often with magnesium particles.

The collisions have more energy too – so more are successful.

So the reaction rate increases with temperature for *two* reasons: there are more collisions, and more of them have enough energy to be successful.

**Why rate increases with surface area** The reaction between the magnesium and acid is much faster when the metal is powdered:

The acid particles can collide only with the magnesium atoms in the outer layer of the metal ribbon.

In the powdered metal, many more atoms are exposed. So the chance of a collision increases.

**Q**

1. Copy and complete: Two particles can react together only if they …… and the …… has enough …… to be ……

2. What is meant by:
   a. a successful collision?
   b. an unsuccessful collision?

3. Reaction between magnesium and acid speeds up when:
   a. the concentration of the acid is doubled. Why?
   b. the temperature is raised. Why? (Give two reasons.)
   c. the acid is stirred. Why?
   d. the metal is ground to a powder. Why?
10.6 Catalysts

What is a catalyst?
You saw that a reaction can be speeded up by increasing the temperature, or the concentration of a reactant, or the surface area of a solid reactant. There is another way to increase the rate of some reactions: use a catalyst.

A catalyst is a substance that speeds up a chemical reaction, but remains chemically unchanged itself.

Example: the decomposition of hydrogen peroxide
Hydrogen peroxide is a colourless liquid that breaks down very slowly to water and oxygen:

\[
\text{hydrogen peroxide} \rightarrow \text{water} + \text{oxygen} \\
2\text{H}_2\text{O}_2 (l) \rightarrow 2\text{H}_2\text{O} (l) + \text{O}_2 (g)
\]

You can show how a catalyst affects the reaction, like this:

1. Pour some hydrogen peroxide into three measuring cylinders. The first one is the control.
2. Add manganese(IV) oxide to the second, and raw liver to the third.
3. Now use a glowing wooden splint to test the cylinders for oxygen. The splint will burst into flame if there is enough oxygen present.

The results

Since hydrogen peroxide breaks down very slowly, there is not enough oxygen to relight the splint.

Manganese(IV) oxide makes the reaction go thousands of times faster. The splint bursts into flame.

Raw liver also speeds it up. The liquid froths as the oxygen bubbles off – and the splint relights.

So manganese(IV) oxide acts as a catalyst for the reaction. If you add more manganese(IV) oxide, the reaction will go even faster.

Something in the raw liver acts as a catalyst too. That ‘something’ is an enzyme called catalase.

What are enzymes?
Enzymes are proteins made by cells, to act as biological catalysts.
Enzymes are found in every living thing. You have thousands of different enzymes inside you. For example catalase speeds up the decomposition of hydrogen peroxide in your cells, before it can harm you. Amylase in your saliva speeds up the breakdown of the starch in your food.

Without enzymes, most of the reactions that take place in your body would be far too slow at body temperature. You would die.
How do catalysts work?
For a reaction to take place, the reacting particles must collide with enough energy for bonds to break and reaction to occur.
When a catalyst is present, the reactants are able to react in a way that requires less energy.
This means that more collisions now have enough energy to be successful. So the reaction speeds up. But the catalyst itself is unchanged.
Note that a catalyst must be chosen to suit the particular reaction. It may not work for other reactions.

Catalysts in the chemical industry
In industry, many reactions need heat. Fuel can be a very big expense.
With a catalyst, a reaction goes faster at a given temperature. So you get the product faster, saving time. Even better, it may go fast enough at a lower temperature – which means a lower fuel bill.
So catalysts are very important in the chemical industry. They are often transition elements or their oxides. Two examples are:
- iron used in the manufacture of ammonia
- vanadium(IV) oxide used in the manufacture of sulfuric acid.

Making use of enzymes
There are thousands of different enzymes, made by living things. We are finding many uses for them.
For example some bacteria make enzymes that catalyse the breakdown of fat, starch, and proteins. The bacteria can be grown in tanks, in factories. The enzymes are removed, and used in biological detergents. In the wash, they help to break down grease, food stains, and blood stains on clothing.
Enzymes work best in conditions like those in the living cells that made them.
- If the temperature gets too high, the enzyme is destroyed or denatured. It loses its shape.
- An enzyme also works best in a specific pH range. You can denature it by adding acid or alkali.

More means faster
The more catalyst you add, the faster the reaction goes.

Q
1 What is a catalyst?
2 Which of these does a catalyst not change?
   a the speed of a reaction
   b the products that form
   c the total amount of each product formed
3 Explain what an enzyme is, and give an example.
4 Why do our bodies need enzymes?
5 Why do our bodies need enzymes?
6 Give two examples of catalysts used in the chemical industry.
7 A box of biological detergent had this instruction on the back: Do not use in a wash above 60 °C. Suggest a reason.
More about enzymes

Mainly from microbes
Enzymes are proteins made by living things, to act as catalysts for their own reactions. So we can obtain enzymes from plants, and animals, and microbes such as bacteria and fungi. In fact we get most from microbes.

Traditional uses for enzymes
We humans have used enzymes for thousands of years. For example ...

- **In making bread**  Bread dough contains yeast (a fungus), and sugar. When the dough is left in a warm place, the yeast cells feed on the sugar to obtain energy. Enzymes in the yeast catalyse the reaction, which is called **fermentation**:

\[
\ce{C_6H_{12}O_6(aq) + \text{enzymes} \rightarrow 2C_2H_5OH(aq) + 2CO_2(g) + \text{energy}}
\]

The carbon dioxide gas makes the dough rise. Later, in the hot oven, the gas expands even more, while the bread sets. So you end up with spongy bread. The heat kills the yeast off.

- **In making yoghurt** To make yoghurt, special bacteria are added to milk. They feed on the lactose (sugar) in it, to obtain energy. Their enzymes catalyse its conversion to lactic acid and other substances, which turn the milk into yoghurt.

Making enzymes the modern way
In making bread and yoghurt, the microbes that make the enzymes are present. But in most modern uses for enzymes, they are not. Instead:

- bacteria and other microbes are grown in tanks, in a rich broth of nutrients; so they multiply fast
- then they are killed off, and their enzymes are separated and purified
- the enzymes are sold to factories.

Anyone home? The tank contains bacteria, busy making enzymes. For example it could be the enzyme amylase, which catalyses the conversion of starch to sugar.
Modern uses of enzymes
Enzymes have many different uses. Here are some common ones:

- **In making soft-centred chocolates**  How do they get the runny centres into chocolates? By using the enzyme **invertase**.
  First they make a paste containing sugars, water, flavouring, colouring, and invertase. Then they dip blobs of it into melted chocolate, which hardens. Inside, the invertase catalyses the breakdown of the sugars to more soluble ones, so the paste goes runny.

Other enzymes are used in a similar way to ‘soften’ food, to make tinned food for infants.

- **In making stone-washed denim**  Once, denim was given a worn look by scrubbing it with pumice stone. Now an enzyme does the job.

- **In making biological detergents**  As you saw on page 137, these contain enzymes to catalyse the breakdown of grease and stains.

- **In DNA testing**  Suppose a criminal leaves tiny traces of skin or blood at a crime scene. The enzyme **polymerase** is used to ‘grow’ the DNA in them, to give enough to identify the criminal.

How do they work?
This shows how an enzyme molecule catalyses the breakdown of a reactant molecule:

First, the two molecules must fit together like jigsaw pieces. (So the reactant molecule must be the right shape, for the enzyme.)

The ‘complex’ that forms makes it easy for the reactant molecule to break down. You do not need to provide energy by heating.

When decomposition is complete the molecules of the product move away. Another molecule of the reactant takes their place.

Enzymes are a much more complex shape than the drawing suggests. Even so, this model gives you a good idea of how they work.

The search for extremophiles
Most of the enzymes we use work around 40°C, and at a pH not far from 7. In other words, in conditions like those in the cells that made them.

But around the world, scientists are searching high and low for microbes that live in very harsh conditions. For example deep under the ice in Antarctica, or at hot vents in the ocean floor, or in acidic lakes around volcanoes. They call these microbes **extremophiles**.

Why do scientists want them? Because the enzymes made by these microbes will work in the same harsh conditions. So they may find a great many uses in industry.
Some reactions need light
Some chemical reactions obtain the energy they need from light. They are called **photochemical reactions**. Examples are photosynthesis, and the reactions that occur in film photography.

**Photosynthesis**
- **Photosynthesis** is the reaction between carbon dioxide and water, in the presence of chlorophyll and sunlight, to produce glucose:

\[
6\text{CO}_2 (g) + 6\text{H}_2\text{O} (l) \xrightarrow{\text{light}} \text{C}_6\text{H}_{12}\text{O}_6 (a\text{q}) + 6\text{O}_2 (g)
\]

- carbon dioxide  water  chlorophyll  glucose  oxygen
- It takes place in plant leaves. Carbon dioxide enters the leaves through tiny holes called **stomata**.
- **Chlorophyll**, the green pigment in leaves, is a **catalyst** for the reaction.
- The water is taken in from the soil, through the plant's roots.
- Sunlight provides the energy for this endothermic reaction.
- The plant then uses the glucose for energy, and to build the cellulose and other substances it needs for growth.

**Changing the rate of the photosynthesis reaction**
Could you change the rate by changing the strength of the light? Let's see.

**The method**  Pondweed is a suitable plant to use for the experiment.

1. Put some pondweed in a beaker containing a very dilute solution of sodium hydrogen carbonate, NaHCO₃. (This compound decomposes, giving off carbon dioxide.) Place a funnel over it.
2. Place a test-tube full of the solution over the funnel, as shown.
3. Place the lamp 50 cm from the beaker. (Look at the arrow above.)
4. Let the pondweed adjust to the conditions for 1 minute. Then count the bubbles of oxygen it gives off, over 1 minute. Repeat twice more to get an average value per minute. Record your results.
5. Repeat step 4, with the lamp placed at 40, 30, 20, and 10 cm from the beaker.

You can then plot a graph for your results.
The results  This graph shows that the number of bubbles per minute *increases* as the lamp is brought closer to the plant. The closer it is, the greater the strength or intensity of the light that reaches the plant. So we can say that the rate of photosynthesis increases as the intensity of the light increases.

That makes sense. Light provides the energy for the reaction. The stronger it is, the more energy it provides. So more molecules of carbon dioxide and water gain enough energy to react.

A photochemical reaction can be speeded up by increasing the intensity of the light. This is true of all photochemical reactions.

The reactions in film photography
Black-and-white film photography relies on a photochemical reaction. The film is covered with a coating of gel that contains tiny grains of silver bromide. Light causes this to break down:

\[ 2\text{AgBr} (s) \rightarrow 2\text{Ag} (s) + \text{Br}_2 (l) \]

It is both a photochemical reaction and a redox reaction.

The silver ions are reduced: \[ 2\text{Ag}^+ + 2e^- \rightarrow 2\text{Ag} \text{  (electron gain)} \]

The bromide ions are oxidised: \[ 2\text{Br}^- \rightarrow \text{Br}_2 + 2e^- \text{  (electron loss)} \]

So how is a photo produced?
1  When you click to take the photo, the camera shutter opens briefly. Light enters and strikes the film. The silver bromide decomposes, giving tiny dark particles of silver. Where brighter light strikes (from brighter parts of the scene), decomposition is faster, giving more silver.

2  Next the film is developed: unreacted silver bromide is washed away, leaving clear areas on the film. The silver remains, giving darker areas.

3  Then the film is printed. In this step, light is shone through the film onto photographic paper, which is also coated with silver bromide. The light passes through the clear areas of the film easily, causing the silver bromide to decompose. But the darker areas block light.

4  The unreacted silver bromide is washed from the paper. This leaves a black-and-white image of the original scene, made of silver particles.
Checkup on Chapter 10

Revision checklist

Core curriculum

Make sure you can ...

- explain what the rate of a reaction means
- describe a way to measure the rate of a reaction that produces a gas, using a gas syringe
- describe a way to measure the rate of a reaction that produces carbon dioxide (a heavy gas), using a balance
- give the correct units for the rate of a given reaction (for example cm³ per minute, or grams per minute)
- work out, from the graph for a reaction:
  - how long the reaction lasted
  - how much product was obtained
  - the average rate of the reaction
  - the rate in any given minute
- give three ways to increase the rate of a reaction
- say which of two reactions was faster, by comparing the slope of their curves on a graph
- explain why there is a risk of explosions in flour mills and coal mines
- explain these terms: catalyst enzyme
- explain why enzymes are important in our bodies
- explain why catalysts are important in industry
- give examples of the use of catalysts, including enzymes

Extended curriculum

Make sure you can also ...

- describe the collision theory
- use the collision theory to explain why the rate of a reaction increases with concentration, temperature (two reasons!) and surface area
- explain how catalysts work
- say how enzymes can be destroyed
- define photochemical reaction and give examples
- say what photosynthesis is, and give the word and chemical equations for it
- name the catalyst for photosynthesis
- explain why a photochemical reaction can be speeded up by increasing the intensity of the light
- give the equation for the photochemical reaction that takes place on black-and-white film and photographic paper
- show that this reaction is also a redox reaction

Questions

Core curriculum

1. The rate of the reaction between magnesium and dilute hydrochloric acid can be measured using this apparatus:

   a. What is the purpose of:
      i. the test-tube?
      ii. the gas syringe?

   b. How would you get the reaction to start?

2. Some magnesium and an excess of dilute hydrochloric acid were reacted together. The volume of hydrogen produced was recorded every minute, as shown in the table:

<table>
<thead>
<tr>
<th>Time / min</th>
<th>0</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
<th>7</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume of hydrogen / cm³</td>
<td>0</td>
<td>14</td>
<td>23</td>
<td>31</td>
<td>38</td>
<td>40</td>
<td>40</td>
<td>40</td>
</tr>
</tbody>
</table>

   a. What does an excess of acid mean?
   b. Plot a graph of the results.
   c. What is the rate of reaction (in cm³ of hydrogen per minute) during:
      i. the first minute?
      ii. the second minute?
      iii. the third minute?
   d. Why does the rate change during the reaction?
   e. How much hydrogen was produced in total?
   f. How long does the reaction last?
   g. What is the average rate of the reaction?
   h. How could you slow down the reaction, while keeping the amounts of reactants unchanged?

3. Suggest a reason for each observation below.
   a. Hydrogen peroxide decomposes much faster in the presence of the enzyme catalase.
   b. The reaction between manganese carbonate and dilute hydrochloric acid speeds up when some concentrated hydrochloric acid is added.
   c. Powdered magnesium is used in fireworks, rather than magnesium ribbon.
4 In two separate experiments, two metals A and B were reacted with an excess of dilute hydrochloric acid. The volume of hydrogen was measured every 10 seconds. These graphs show the results:

![Graph showing volume of hydrogen vs. time for metals A and B.]

a i Which piece of apparatus can be used to measure the volume of hydrogen produced?

b ii What other measuring equipment is needed?

c Which metal, A or B, reacts faster with hydrochloric acid? Give your evidence.

d Sketch and label the curves that will be obtained for metal B if:

i more concentrated acid is used (curve X)

ii the reaction is carried out at a lower temperature (curve Y).

5 Copper(II) oxide catalyses the decomposition of hydrogen peroxide. 0.5 g of the oxide was added to a flask containing 100 cm³ of hydrogen peroxide solution. A gas was released. It was collected, and its volume noted every 10 seconds. This table shows the results:

<table>
<thead>
<tr>
<th>Time / s</th>
<th>0</th>
<th>10</th>
<th>20</th>
<th>30</th>
<th>40</th>
<th>50</th>
<th>60</th>
<th>70</th>
<th>80</th>
<th>90</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume / cm³</td>
<td>0</td>
<td>18</td>
<td>30</td>
<td>40</td>
<td>48</td>
<td>53</td>
<td>57</td>
<td>58</td>
<td>58</td>
<td>58</td>
</tr>
</tbody>
</table>

a What is a catalyst?

b Draw a diagram of suitable apparatus for this experiment.

c Name the gas that is formed.

d Write a balanced equation for the decomposition of hydrogen peroxide.

e Plot a graph of the volume of gas (vertical axis) against time (horizontal axis).

f Describe how rate changes during the reaction.

g What happens to the concentration of hydrogen peroxide as the reaction proceeds?

h What chemicals are present in the flask after 90 seconds?

i What mass of copper(II) oxide would be left in the flask at the end of the reaction?

j Sketch on your graph the curve that might be obtained for 1.0 g of copper(II) oxide.

k Name one other substance that catalyses this decomposition.

Extended curriculum

6 Marble chips (lumps of calcium carbonate) react with hydrochloric acid as follows:

\[
\text{CaCO}_3 (s) + 2\text{HCl (aq)} \rightarrow \text{CaCl}_2 (aq) + \text{CO}_2 (g) + \text{H}_2\text{O (l)}
\]

a What gas is released during this reaction?

b Describe a laboratory method that could be used to investigate the rate of the reaction.

c How will this affect the rate of the reaction?

i increasing the temperature

ii adding water to the acid

d Explain each of the effects in c in terms of collisions between reacting particles.

e If the lumps of marble are crushed first, will the reaction rate change? Explain your answer.

7 Zinc and iodine solution react like this:

\[
\text{Zn (s)} + \text{I}_2 (aq) \rightarrow \text{ZnI}_2 (aq)
\]

The rate of reaction can be followed by measuring the mass of zinc metal at regular intervals, until all the iodine has been used up.

a What will happen to the mass of the zinc, as the reaction proceeds?

b Which reactant is in excess? Explain your choice.

c The reaction rate slows down with time. Why?

d Sketch a graph showing the mass of zinc on the y axis, and time on the x axis.

e How will the graph change if the temperature of the iodine solution is increased by 10 °C?

f Explain your answer to e using the idea of collisions between particles.

8 Some pondweed is placed as shown:

![Diagram showing a test-tube with gas collects, bubbles, pond weed, water containing dissolved carbon dioxide.]

a i Name the gas that collects in the test tube.

ii What other product is produced?

b This experiment must be carried out in the light. Why?

b Using the apparatus above, suggest a method by which the rate of reaction could be found.

d What would be the effect of bringing a lamp close to the beaker? Explain your answer.