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S3 New Curriculum chemistry

Theme: Reactants and products

Chapter 6 – Chemical reactions

Chemical reactions

A chemical reaction is a process in which one or more chemical substances called reactants combine or breakdown to form new substances called products. Substances are either chemicals elements or compounds.

Rate of a chemical reaction:

The **rate of a chemical reaction** refers to how quickly reactants are converted into products.

Ways of measuring the rate of reaction

Scientists use various methods to measure reaction rates depending on the type of reaction. Here are some common techniques:

1. Measuring Change in Mass

- If a reaction produces a gas, the mass of the system will decrease over time.
- A **balance (scale)** can be used to record mass loss.
- Example: **Decomposition of calcium carbonate (CaCO_3) producing CO_2 gas.**

2. Measuring Volume of Gas Produced

- A **gas syringe** or an **inverted measuring cylinder** in water can measure gas volume over time.
- Useful for reactions that generate gases like hydrogen (H_2) or carbon dioxide (CO_2).
- Example: **Reaction of magnesium (Mg) with hydrochloric acid (HCl) producing hydrogen gas.**

3. Monitoring Color Change (Spectrophotometry)

- If the reaction involves a color change, a **spectrophotometer** measures absorbance changes.
- Example: **Reaction of iodine with starch forming a blue-black color.**

4. Tracking pH Changes

- If a reaction produces or consumes acidic or basic substances, a **pH meter** can monitor pH over time.
- Example: **Neutralization of acids and bases.**

5. Measuring Change in Conductivity

- Conductivity changes if ions are formed or consumed during a reaction.
- A **conductivity meter** can track this change.
- Example: **Dissolution of salts in water affecting ionic concentration.**

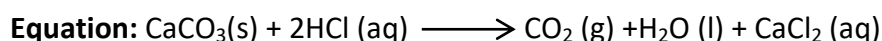
6. Measuring Precipitate Formation

- Some reactions form insoluble solids (precipitates).
- **Cloudiness/turbidity** of the solution can be measured.
- Example: **Reaction acid with sodium thiosulphate to form sulphur**

7. Tracking Temperature Changes

- Exothermic or endothermic reactions alter the temperature.
- A **thermometer** or **calorimeter** records temperature change.
- Example: **Combustion reactions or dissolution of salts in water.**

Experiment to determine the rate of reaction of calcium carbonate with dilute hydrochloric acid by measuring the rate production carbon dioxide



Setup

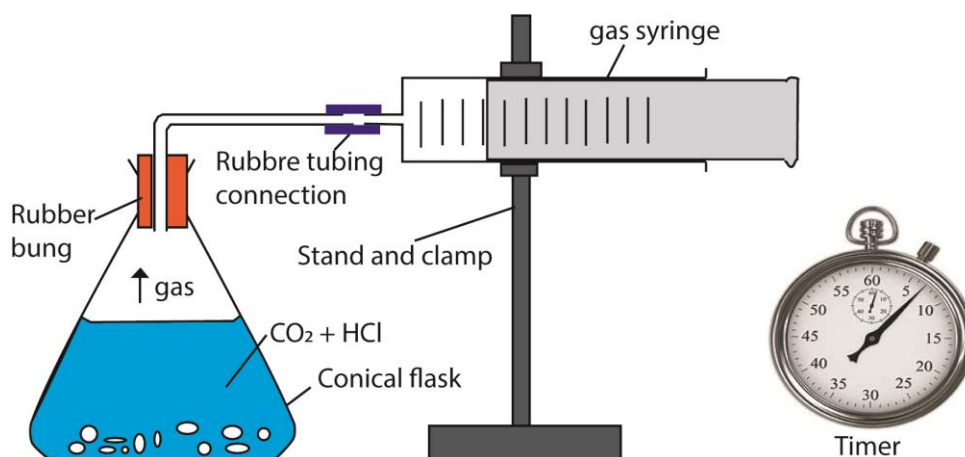


Figure 1

Procedure

- Measure 50cm³ of dilute hydrochloric acid using a measuring cylinder into the conical flask.
- Add 10g of calcium carbonate to the acid close the rubber bung and immediately start a stop clock
- Read and record the volume of carbon dioxide gas produced in the syringe every five second and record your results in the table below

Time in second	5	10	15	20	25	30	35	40
Volume of CO ₂ produced (cm ₃)								

- Plot a graph showing how the volume of carbon dioxide changes with time.
- The rate of reaction at any time is determined from the slope of the gradient at that particular instance.

Figure 2 (a) Shows how the rate of reaction at time t is determined by drawing a tangent at that time while (b) shows curves for the reactions with higher rates (or fast reaction) and for slow reactions.

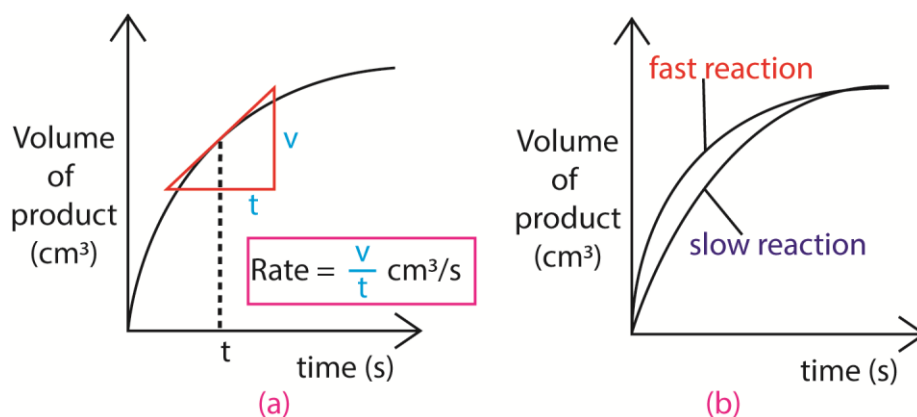


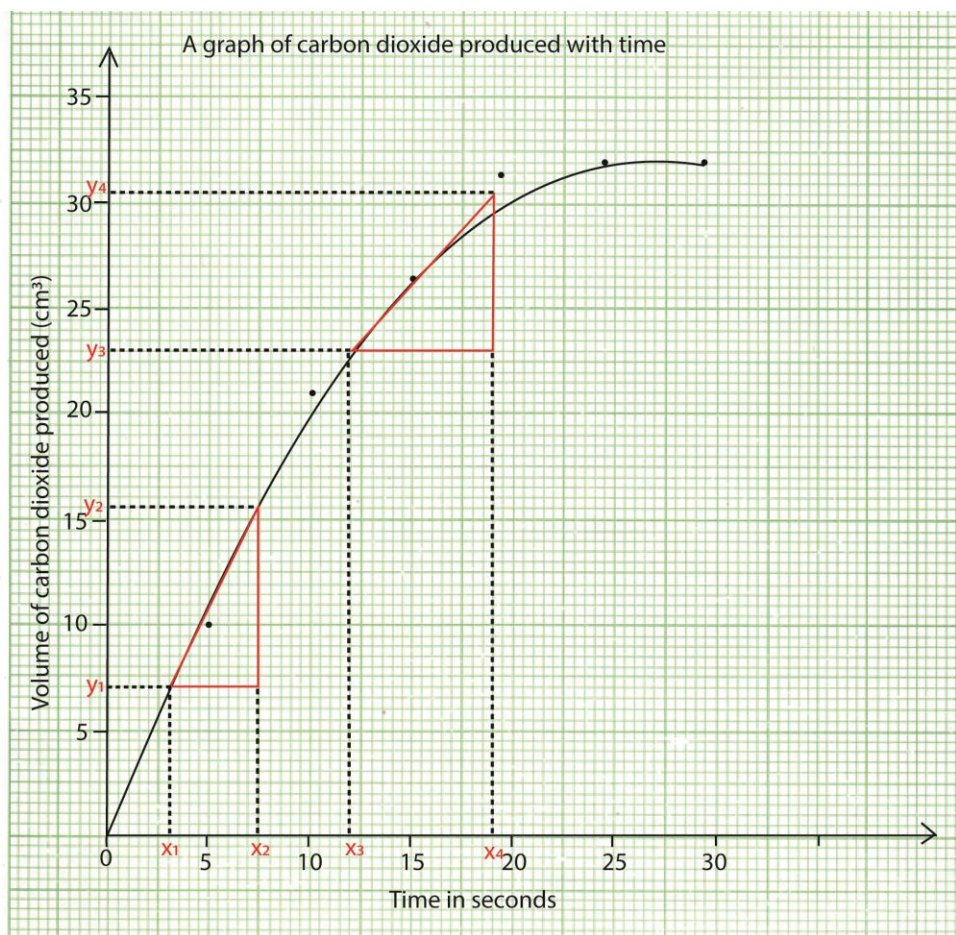
Fig. 2

Example 1

The table below shows the volume of carbon dioxide liberated with time when dilute hydrochloric acid was reacted with calcium carbonate at 25⁰C.

Time in second	0	5	10	15	20	25	30
Volume of CO ₂ produced (cm ₃)	0	10	21	26.5	31.5	32	32

- Plot the graph of volume of hydrogen evolved (vertical) against time (horizontal axis)



(ii) Determine the rate of reaction at 5 seconds.

$$\text{Rate at 5 seconds} = \frac{y_2 - y_1}{x_2 - x_1} = \frac{15.5 - 7.0}{7.5 - 3} = \frac{8.5}{4.5} = 1.9 \text{ cm}^3/\text{s}$$

(iii) Determine the rate of reaction at 20 seconds

$$\text{Rate at 20 seconds} = \frac{y_4 - y_3}{x_4 - x_3} = \frac{30.5 - 23.0}{19 - 12} = \frac{7.5}{7.0} = 1.1 \text{ cm}^3/\text{s}$$

(iv) Explain the difference between the rates at 12 seconds and 20 seconds

The rate is higher at 5 seconds than at 20 seconds because the concentrations of the reactants at 5 seconds are higher than those at 20 seconds

(v) Explain why the rate does not change between 25 and 30 seconds

Because one of the reagents is used up

Factors that affect a chemical reaction

There are about four factors that affect a chemical change/reaction. These are:

1. A catalyst
2. Temperature
3. Concentration
4. Pressure
5. Surface area of the reactants

1. Concentration:

Increase in concentration increases the rate of reaction due to increase in the frequency of molecular collision.

Experiment: Investigating the Effect of Concentration on Reaction Rate

Aim: To determine how changing the concentration of **hydrochloric acid (HCl)** affects the rate of reaction with **magnesium (Mg) strips**, measured by the time taken for a given length of magnesium strip to react with equal volume of acid of different concentration.

Materials Needed:

- Hydrochloric acid (HCl) solutions of different concentrations (e.g., 0.5 M, 1 M, 2 M)
- Magnesium strips (Mg)
- Measuring cylinder
- Stopwatch
- Beaker
- Gas syringe (or an inverted measuring cylinder in water to collect gas)
- Safety goggles and gloves

Procedure:

1. **Prepare Different HCl Solutions** – Pour 50 mL of HCl solution into separate beakers with varying concentrations (0.5 M, 1 M, 2 M, 3M).
2. **Add Magnesium Strip** – Drop a magnesium strip into the first beaker containing 0.5 M HCl.
3. **Measure Time and Gas Volume** – Start the stopwatch and record the time taken for the reaction to **stop fizzing** (when hydrogen gas stops forming).
4. **Repeat for Different Concentrations** – Perform the experiment using magnesium strips of equal length with 1 M and 2 M HCl, recording reaction time and gas volume produced.
5. **Compare Results** – Observe how higher concentrations of HCl cause faster reactions and **more rapid gas production**.

Sample results table

HCl Concentration (M)	0.5	1	2	2
Time for Reaction to Stop (s)	90	60	42	25

Observation

The **3 M HCl solution** reacts the fastest, completing in **25 seconds**, while **0.5 M HCl** takes **90 seconds**.

Conclusion:

Higher concentration of HCl → Faster reaction, i.e. Increasing HCl concentration results in **more frequent collisions** between acid molecules and magnesium atoms, speeding up the reaction.

Experiment: Investigating the Effect of Concentration on Reaction Rate

Aim: To determine how the concentration of **sodium thiosulphate** affects the rate of reaction with **hydrochloric acid**, measured by how quickly the solution turns **opaque**.

Materials Needed:

- Sodium thiosulphate solution (different concentrations: **0.1 M, 0.2 M, 0.5 M, 1 M**)
- Hydrochloric acid (HCl, constant concentration)
- Conical flask
- Stopwatch
- White paper with a black "X" mark
- Measuring cylinders
- Safety goggles and gloves

Procedure:

- Prepare different sodium thiosulphate solutions** – Pour 50 mL into separate 100 cm³ flasks with varying concentrations (**0.1 M, 0.2 M, 0.5 M, 1 M**).
- Place flask containing 0.1M sodium thiosulphate solution on the marked paper** – The black "X" helps measure how quickly the solution becomes cloudy.
- Add hydrochloric acid** – Pour **10 mL of HCl** into the flask and immediately start the stopwatch.
- Observe and record time** – Stir the mixture and note the time when the "X" **completely disappears** due to the formation of sulfur precipitate.
- Repeat the experiment** – Perform the reaction for each concentration of sodium thiosulphate.

Sample Results Table:

Sodium thiosulphate solution	0.1M	0.2M	0.5M	1M
Time taken for mark "X" to disappear (seconds)	90	65	40	24

Observation

1 M sodium thiosulphate reacted in **25 seconds** (Shorter time), while **0.1 M** took **90 seconds** (longer time)

Conclusion:

- **Higher concentration → Faster reaction**
- **More concentrated solutions lead to a quicker disappearance of the "X"** because more particles are available for reaction, increasing the frequency of collisions.

2. Temperature:

Increasing the temperature increases reaction rates because of the disproportionately large **increase** in the number of high energy collisions. It is only these collisions (possessing at least the activation energy for the **reaction**) which result in a **reaction**.

Experiment: Investigating the Effect of Temperature on Reaction Rate

Aim: To determine how changing the temperature of **hydrochloric acid** affects the rate of its reaction with **magnesium**, measured by the volume of hydrogen gas produced.

Materials Needed:

- Hydrochloric acid (1 M)
- Magnesium strips (same size for consistency)
- Beaker (100 mL)
- Water bath or hot plate (for heating solutions)
- Thermometer
- Gas syringe or measuring cylinder (for collecting gas)
- Stopwatch
- Safety goggles and gloves

Procedure:

- Pour **50 mL** of the acid into 100ml beaker.
- Warm the acid to **30°C**
- Drop a magnesium strip into the beaker.
- Start stopwatch and begin timing the reaction immediately.
- (v)** Measure the time taken for the reaction to stop.
- (vi)** Repeat the experiment with the acid warmed 35°C, 45°C and 60°C.

Sample result Table

Temperature of the acid (°C)	30	35	45	60
Time taken for the reaction to stop (seconds)	120	80	50	30

Observations

The **60°C acid** reacted the fastest, completing in **30 seconds**, while the **30°C acid** took **120 seconds**.

Conclusion:

Higher temperature → Faster reaction because at higher temperature the reaction took a shorter time. At higher temperature, particles move **faster**, increasing the **frequency of collisions** and **energy for successful reactions**.

3. Surface area:

Increasing the surface area of a solid reactant exposes more of its particles to attack. This results in an **increased** chance of collisions between reactant particles, so there **are** more collisions in any given time and the **rate of reaction increases**. The surface of a solid can be increased by crushing a substance into a powder or by reducing its particle size.

Experiment: Investigating the Effect of Surface Area on Reaction Rate

Aim: To determine how the surface area of **calcium carbonate (CaCO₃)** affects the rate of reaction with **hydrochloric acid (HCl)**, measured by the time taken for equal mass of chips and powder to react with equal volumes of the acid.

Materials Needed:

- Calcium carbonate in different forms:
 - **Large marble chips** (low surface area)
 - **Powdered calcium carbonate** (high surface area)
- Hydrochloric acid (1 M)
- Measuring cylinder (50 mL)
- Conical flask
- Stopwatch
- Safety goggles and gloves

Procedure:

- (a) **Measure hydrochloric acid** – Pour **50 mL of HCl** into a conical flask.
- (b) **Add marble chips** – Drop **5 g** of large calcium carbonate chips into the acid.
- (c) **Start stopwatch** – Immediately begin timing the reaction.
- (d) **Measure time taken for the reaction to complete.**
- (e) **Repeat for powder** – Perform the experiment keeping all other conditions constant.

Sample Results Table:

Calcium Carbonate form	Large Chips	Powder
Time for Reaction to Stop (seconds)	120	42

Observations

The powdered form reacts the fastest (**40s**), while the marble chips react the slowest (**115s**).

Conclusion:

- **Smaller particle size (or large surface area) → Faster reaction**
- **Higher surface area the shorter the time taken for the reaction** because more particles are exposed for reaction.

4. Pressure:

This affects gaseous reactions since gases unlike solids are compressible. Increasing pressure on gases brings reactant particles close to each other increasing the frequency of collision and hence the rate of reaction. Pressure can be increased by decreasing the volume of the container. For example, in the Haber process, a large yield of ammonia is obtained from high pressure as per the following reaction.

5. **Catalyst:** increases the rate of reaction by lowering activation energy or energy barrier to the reaction.

By definition

A catalyst is substance that increases the rate of a chemical reaction without itself undergoing any permanent chemical change.

To investigate the effect of a catalyst on the production of oxygen

You are provided with three test tubes labelled A, B, C; to test tube

A put hydrogen peroxide

B put manganese dioxide

C put both manganese dioxide first followed with hydrogen peroxide

Allow the experiment to stand for 10 minutes.

Insert a glowing splint in test tube A, B C respectively

Record your observation and conclusion.

Observations

In **A** oxygen was given off after a long period. The glowing splint kept on glowing.

In **B** no oxygen was given off.

In **C** oxygen was given off after a short time. The glowing splint busted into a flame

Conclusion:

The manganese dioxide speeds up the production of oxygen from hydrogen peroxide.

Reversible reactions

Reversible reactions are chemical reactions where the reactants can form products, but the products can also revert back to the original reactants under certain conditions.

A reversible reaction where A and B react to form C and D can be represented as follows

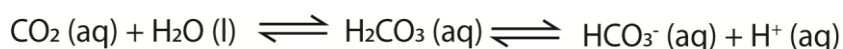


At equilibrium the rate of the formation of the product is equal to the rate of formation of the reactants.

Chemical equilibrium

Chemical equilibrium is a state in which the rate of the forward reaction equals the rate of the backward reaction. In other words, there is no net change in the concentrations of the reactants and the product. This kind of equilibrium is also called dynamic equilibrium.

An example of reversible reactions are Dissolution of carbon dioxide in soda at high pressure.



At high pressure carbon dioxide dissolves in water to form carbonic acid that hydrolyse in water to form hydrogen carbonate ion and hydrogen ions. When the bottle is opened, pressure decreases and this reaction reverses.

Le Chatelier's Principle

This principle states that when a system at equilibrium is disturbed, it **adjusts to counteract the change** and restore equilibrium.

Factors that affect reversible reactions

According to **Le Chatelier's Principle**, several factors influence reversible reactions and can shift the equilibrium between reactants and products.

1. Concentration of Reactants and Products

- **Increasing reactant concentration** → Shifts equilibrium forward, producing more products.
- **Increasing product concentration** → Shifts equilibrium backward, favoring reactants.
- Example: In the **Haber process** ($\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$), adding more nitrogen or hydrogen increases ammonia production.

2. Temperature

- **For exothermic reactions** (heat is released): Increasing temperature shifts equilibrium **backward**, favoring reactants.
- **For endothermic reactions** (heat is absorbed): Increasing temperature shifts equilibrium **forward**, producing more products.
- Example: **Decomposition of calcium carbonate** ($\text{CaCO}_3 \rightleftharpoons \text{CaO} + \text{CO}_2$) is endothermic, so heating favors decomposition.

3. Pressure (For Gaseous Reactions)

- **Increasing pressure** → Shifts equilibrium toward the side with fewer gas molecules.
- **Decreasing pressure** → Shifts equilibrium toward the side with more gas molecules.
- **Example:** In the Haber process, increasing pressure favors ammonia formation because fewer gas molecules exist on the product side.

4. Presence of a Catalyst

- A **catalyst** speeds up both forward and backward reactions **without shifting equilibrium**.
- It **reduces activation energy**, allowing equilibrium to be reached faster.
- Example: Iron is used as a catalyst in the Haber process but does not change the equilibrium position.

5. Nature of the Reaction

- Some reactions naturally **favor one direction** based on their energy requirements or stability of products.
- Example: **Weak acid dissociation** ($\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$) favors mostly reactants because acetic acid only partially dissociates.

Application of reversible reaction

Manufacture of ammonia

By reaction nitrogen and hydrogen



Conditions for production of ammonia

1. Pressure:

Production of ammonia is accompanied with a decrease in the number of moles of a gas and as such high yield of ammonia is favored by high pressure.

2. Temperature

Production of ammonia produces heat; high yields are thus expected at low temperatures.

However, at low temperature, the reaction is slow that a compromise temperature of 450^o-500^oC is used.

3. A catalyst is iron

Uses of ammonia

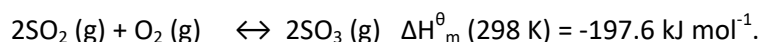
- Disinfectant
- For preparation of liquid soap
- Fertilizers

Note that prolonged use of ammonia nitrate fertilizer may make the soil acidic because ammonium nitrate in soil reacts with water forming ammonium hydroxide and nitric acid plants absorb ammonium ions leaving nitric acid in solution.



The contact process

In sulphur dioxide is reacted with oxygen in presence of Vanadium (V) oxide (catalyst) to form sulphur trioxide. The conversion of sulphur dioxide and oxygen into sulphur trioxide (sulphur (VI) oxide) is an exothermic reaction with a decrease in gas volume.



Condition for production of sulphur trioxide

Temperature: 420^oC

Catalyst : vanadium (V) Oxide

Pressure:

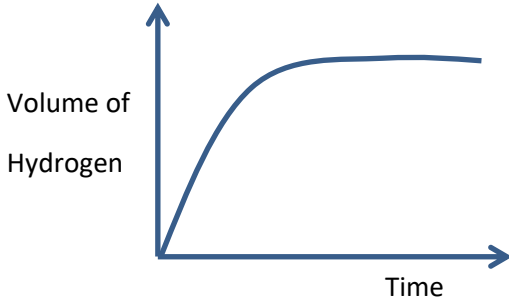
An increase in pressure would increase the yield of the reaction because there are fewer number of moles of gas on the right hand side. However, a pressure of 1 atm is enough to obtain a very good yield. Thus it would be a waste of money and resources to increase the pressure

Revision exercise

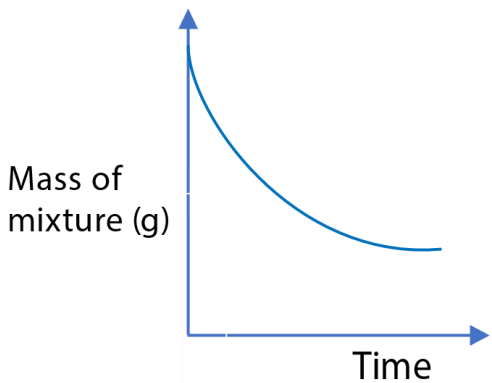
1.		<p>Curve Y in diagram shows the results that were obtained during the investigation of the rate of the reaction between iron and dilute hydrochloric acid under normal conditions. Curve X and Z were obtained when some conditions of the experiment were changed.</p>		
	(a)	(i)	List three condition that were changed to obtain curve X	(3marks)
		(ii)	State what point M represent	(½ mark)
	(b)		Some conditions you have listed in (a)(i) were changed to obtained curve Z.	
		(i)	State the condition changed	(01mark)
		(ii)	Give a reason for your answer	
2.			<p>Sodium thiosulphate reacts with hydrochloric acid according to the following equation</p> $\text{S}_2\text{O}_3^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) \longrightarrow \text{H}_2\text{O}(\text{l}) + \text{SO}_2(\text{g}) + \text{S}(\text{s})$	
	(a)		State what would be observed if dilute hydrochloric acid is added to sodium thiosulphate solution	(½ mark)
	(b)		The rate of reaction with dilute acids is affected by the concentration of sodium thiosulphate.	
		(i)	State one other factor other than concentration that can affect the rate of reaction	(½ mark)
		(ii)	Briefly explain the effect of the factor you have state in (b)(i) on the rate of reaction.	(02mark)
		(iii)	Describe an experiment can be carried out in the laboratory to show the effect of the factor on have stated in (b)(i) on the rate of reaction. Diagram not required)	(6 ½ marks)
3.			State and explain the effect of each of the following conditions on the rate of chemical reaction.	
	(a)		Particle size	(04marks)

	(b)		Concentration of reactants	(05 msrks)																		
	(c)		Temperature	(06marks)																		
4.			<p>The table shows the variation in the concentration of sodium thiosulphate with time.</p> <table border="1"> <tr> <td>Time (s)</td> <td>200</td> <td>100</td> <td>40</td> <td>20</td> <td>10</td> </tr> <tr> <td>Concentration of thiosulphate (mol dm^{-3})</td> <td>0.05</td> <td>0.09</td> <td>0.15</td> <td>0.20</td> <td>0.25</td> </tr> <tr> <td>$\frac{1}{\text{concentration}}$ of thiosulphate ($\text{mol}^{-1} \text{m}^{-3}$)</td> <td></td> <td></td> <td></td> <td></td> <td></td> </tr> </table>	Time (s)	200	100	40	20	10	Concentration of thiosulphate (mol dm^{-3})	0.05	0.09	0.15	0.20	0.25	$\frac{1}{\text{concentration}}$ of thiosulphate ($\text{mol}^{-1} \text{m}^{-3}$)						
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Concentration of thiosulphate (mol dm^{-3})	0.05	0.09	0.15	0.20	0.25																	
$\frac{1}{\text{concentration}}$ of thiosulphate ($\text{mol}^{-1} \text{m}^{-3}$)																						
		(i)	Determine the values of $\frac{1}{\text{concentration}}$ of thiosulphate, copy the table and enter your answer in the space provided in the table.	(01mark)																		
		(ii)	Plot a graph of $\frac{1}{\text{concentration}}$ of thiosulphate (vertical axis) against time (horizontal axis).	(03marks)																		
		(iii)	State any conclusion that can be drawn from the shape of the graph.	(1 ½ marks)																		
5.	(a)	(i)	Write an equation for the reaction between dilute nitric acid and calcium carbonate	(1 ½ mark)																		
		(ii)	State how temperature can affect the rate of reaction in (a)(i) above	(01mark)																		
		(iii)	Give a reason why a large surface area of calcium carbonate can speed up rate of reaction in (a)(i)	(2 marks)																		
	(b)		Magnesium can react with hydrochloric acid to form hydrogen. State the conditions and write equation for the reaction	(2 ½ marks)																		
	(c)		<p>The table below shows the volume of hydrogen evolved when various lengths of magnesium ribbon were reacted with fixed volume of hydrochloric acid</p> <table border="1"> <tr> <td>Length of ribbon (cm)</td> <td>1.0</td> <td>2.0</td> <td>3.0</td> <td>5.0</td> <td>6.0</td> </tr> <tr> <td>Volume of hydrogen (cm^3/min)</td> <td>2.2</td> <td>3.6</td> <td>5.2</td> <td>9.2</td> <td>10.8</td> </tr> </table>	Length of ribbon (cm)	1.0	2.0	3.0	5.0	6.0	Volume of hydrogen (cm^3/min)	2.2	3.6	5.2	9.2	10.8							
Length of ribbon (cm)	1.0	2.0	3.0	5.0	6.0																	
Volume of hydrogen (cm^3/min)	2.2	3.6	5.2	9.2	10.8																	
		(i)	Plot a graph of volume of hydrogen (vertical axis) against length of magnesium ribbon (horizontal axis)	(5marks)																		

	(ii)	Explain the shape of the graph you have drawn.	(2marks)																		
	(iii)	Using your graph, determine the rate of reaction if 4.0cm of magnesium ribbon was used	(01mark)																		
6.		Oxygen is formed from hydrogen peroxide in the presence of Manganese dioxide according to the following equation $2\text{H}_2\text{O}_2(\text{aq}) \xrightarrow{\text{MnO}_2} 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$																			
	(a)	In an experiment, a certain volume of hydrogen peroxide was used to prepare oxygen at room temperature. With the aid of a suitable diagram describe how the following can be determined.																			
	(i)	The volume of oxygen evolved	(5 ½ marks)																		
	(ii)	The rate of evolution of oxygen	(03marks)																		
	(b)	In another experiment one half of volume of hydrogen peroxide in (a) was diluted with equal volume of water. On the same axes draw graphs to show the variation of volume of oxygen with time in (a) and (b).	(03marks)																		
	(c)	Oxygen produced from 200cm ³ of 0.5M hydrogen peroxide solution was reacted completely with magnesium. Calculate the mass of magnesium that reacted	(3 ½ marks)																		
7.	(a)	Define the term rate of reaction	(02marks)																		
	(b)	The table below shows variation of hydrogen evolved with time when dilute hydrochloric acid was added to excess zinc. <table border="1" style="margin-left: auto; margin-right: auto;"> <tbody> <tr> <td>Volume of hydrogen (cm³)</td> <td>0</td> <td>20</td> <td>35</td> <td>46</td> <td>56</td> <td>72</td> <td>79</td> <td>79</td> </tr> <tr> <td>Time (s)</td> <td>0</td> <td>10</td> <td>20</td> <td>30</td> <td>40</td> <td>60</td> <td>80</td> <td>90</td> </tr> </tbody> </table> Plot a graph of volume hydrogen evolved against time	Volume of hydrogen (cm ³)	0	20	35	46	56	72	79	79	Time (s)	0	10	20	30	40	60	80	90	(05mrks)
Volume of hydrogen (cm ³)	0	20	35	46	56	72	79	79													
Time (s)	0	10	20	30	40	60	80	90													
	(c)	Using the graph determine the time taken to collect 60cm ³ of hydrogen gas	(01mark)																		
	(d) (i)	Draw tangents on your graph at points when the time is 20 and 60 seconds and determine the gradient of each tangent.	(04marks)																		

	(ii)	Compare the rate of reaction at 20 and 60seconds. Explain your answer	(03marks)
8.		<p>The graph in figure below shows the variation of volume of hydrogen evolved with time when excess magnesium was added to 100cm³ of 1.0M sulphuric at room temperature.</p> 	
	(a)	Calculate the number of moles of hydrogen ions that is contained in 100cm ³ of	
	(i)	A 0.5M H ₂ SO ₄	(01mark)
	(ii)	A 1.0M H ₂ SO ₄	(01mark)
	(b) (i)	Sketch on the same axes of the graph in figure above the graph that would be obtained if the same mass of magnesium was added to 100cm ³ of a 0.5M sulphuric acid at room temperature.	(01mark)
	(ii)	Mark on the graph the times the two reactions have reached completion.	(01mark)
	(iii)	Compare the time the reaction took to reach completion when 0.5M sulphuric acid was used to that when 1.0M sulphuric acid was used	(01mark)
9.	(a)	<p>Describe an experiment to show how surface area can affect the rate of reaction between calcium carbonate and 2M hydrochloric acid. Your answer must include:</p> <ul style="list-style-type: none"> • A labelled diagram of apparatus • Sketch of expected graph • Mention how the graph can be used to reach a conclusion. 	11 ½
	(b)	Briefly explain why, when 4M hydrochloric acid was used instead of the 2M acid, the rate of reaction was faster. Explain this observation.	(2½ marks)
	(c)	State one other factor than mentioned above that can affect the rate of reaction between hydrochloric acid and calcium carbonate	(01mmark)

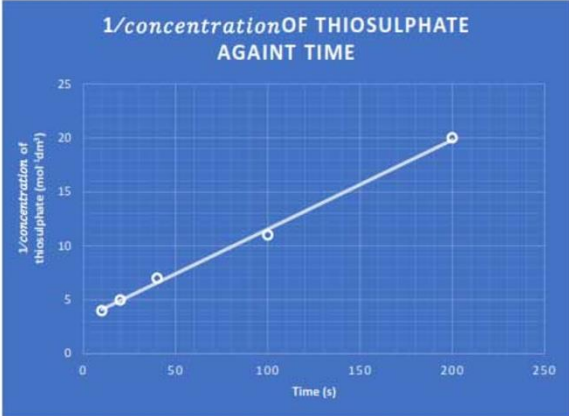
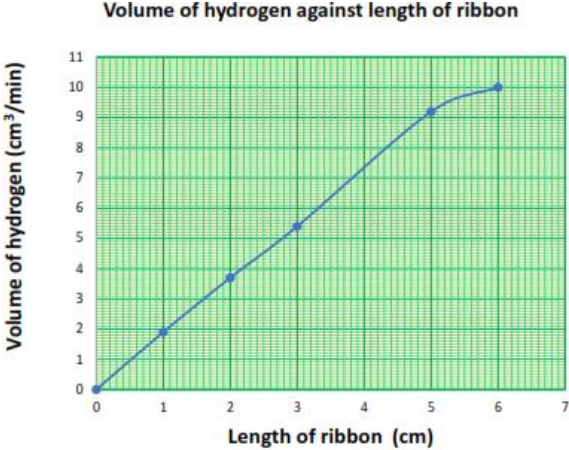
10.	(a)	What is meant by the rate of reaction?	(02marks)																		
	(b)	State how the following factors affect the rate of a chemical reaction:																			
	(i)	Temperature																			
	(ii)	Surface area of the reactants																			
	(c)	<p>The table below shows the volume of hydrogen collected at various time intervals when magnesium was reacted with a 2M hydrochloric acid</p> <table border="1" style="margin-left: auto; margin-right: auto;"> <tr> <td>Time (s)</td> <td>0</td> <td>1</td> <td>2</td> <td>3</td> <td>4</td> <td>5</td> <td>6</td> <td>7</td> </tr> <tr> <td>Volume of hydrogen collected (cm³)</td> <td>0</td> <td>25</td> <td>45</td> <td>60</td> <td>70</td> <td>75</td> <td>77</td> <td>77</td> </tr> </table>	Time (s)	0	1	2	3	4	5	6	7	Volume of hydrogen collected (cm ³)	0	25	45	60	70	75	77	77	
Time (s)	0	1	2	3	4	5	6	7													
Volume of hydrogen collected (cm ³)	0	25	45	60	70	75	77	77													
	(i)	Plot a graph of volume of hydrogen versus time	(04 marks)																		
	(ii)	Determine the rate of reaction at 3 seconds	(02marks)																		
	(iii)	Determine the volume of hydrogen evolved at 3.5 seconds	(01mark)																		
	(d)	State how the rate of reaction at 3 seconds would be affected if a 1M hydrochloric acid was used	(02marks)																		
11.		When a certain volume of 0.1M hydrochloric acid was reacted at room temperature with excess iron fillings, 120cm ³ of the gas were produced.																			
	(a)	Draw a labelled diagram to show how the rate of reaction was determined																			
	(b)	Write equation for the reaction that took place																			
	(c)	Calculate the																			
	(i)	Volume of hydrochloric acid required to produce 120cm ³ of the gas																			
	(ii)	The mass of iron filling that reacted																			
	(d)	Draw a sketch graph of the volume of the gas against time																			
	(e)	State how the rate of reaction would change if the reaction was carried out at a temperature above room temperature																			
12.	(a)	State the factors that can affect the rate of chemical reaction.																			
	(b)	A mixture of a known mass of magnesium and a certain volume of 2M hydrochloric acid were put in a conical flask and the mass of the mixture was recorded at various intervals. The																			

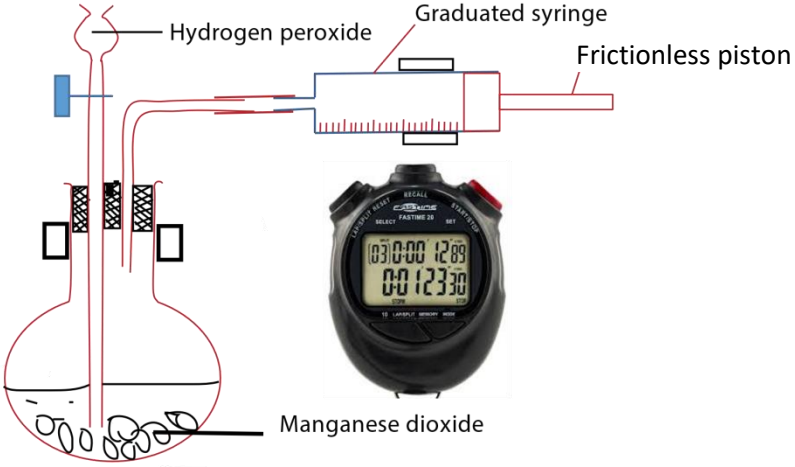
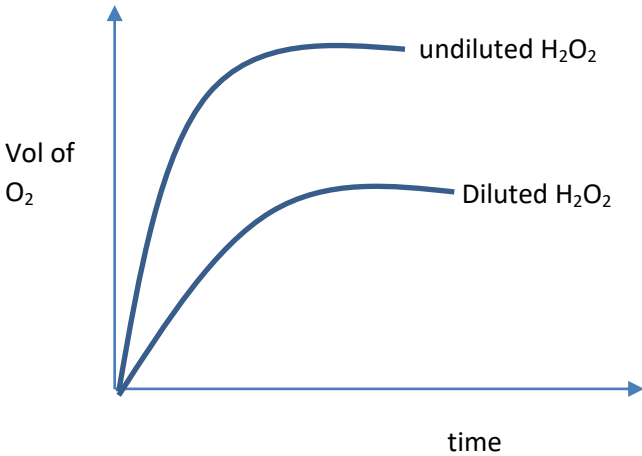
		<p>results of the experiment are shown in the graph below.</p>  <p>Mass of mixture (g)</p> <p>Time</p> <p>On the same axes, draw a graph that would be obtained when same mass of magnesium was reacted with the same volume of 1M hydrochloric acid.</p>																			
	(c)	5.0g of calcium carbonate was reacted with 20cm ³ of 2M hydrochloric acid.																			
		(i) Write equation for the reaction between hydrochloric acid and calcium carbonate																			
		(ii) The mass of calcium carbonate that was left. (Ca = 40, C = 12, O =16)																			
13.	(a)	(i) What is the rate of reaction?																			
		(ii) How does the particle size affect rate of reaction? Explain your answer																			
	(b)	<p>The table of results shows the time taken for sulphur to form when various concentrations of sodium thiosulphate were used.</p> <table border="1" data-bbox="399 1433 1157 1702"> <tr> <td>Concentration of S₂O₃²⁻ (M)</td> <td>0.2</td> <td>0.6</td> <td>0.8</td> <td>1.2</td> <td>1.6</td> </tr> <tr> <td>Time for sulphur to form (s)</td> <td>60</td> <td>20</td> <td>15</td> <td>10</td> <td>7.5</td> </tr> <tr> <td>1/t (s⁻¹)</td> <td>0.017</td> <td>0.05</td> <td>0.07</td> <td>0.10</td> <td>0.13</td> </tr> </table> <p>Plot a graph of 1/t on vertical axis against concentration of thiosulphate.</p>	Concentration of S ₂ O ₃ ²⁻ (M)	0.2	0.6	0.8	1.2	1.6	Time for sulphur to form (s)	60	20	15	10	7.5	1/t (s ⁻¹)	0.017	0.05	0.07	0.10	0.13	
Concentration of S ₂ O ₃ ²⁻ (M)	0.2	0.6	0.8	1.2	1.6																
Time for sulphur to form (s)	60	20	15	10	7.5																
1/t (s ⁻¹)	0.017	0.05	0.07	0.10	0.13																
	(c)	(i) Explain the relationship between the rate of the reaction and 1/t.																			
		(ii) Deduce from the graph, how the rate of reaction varies with																			

			the concentration of thiosulphate.																	
	(d)		Name one reagent that you would use to test for sulphur dioxide and state what would be observed if the reagent was used.																	
14.			In an experiment to determine the rate of reaction between zinc and sulphuric acid, dilute sulphuric acid was reacted with zinc granules to which some copper II sulphate solution was added. The volume of hydrogen gas evolved at various <table border="1" data-bbox="405 568 1158 689"> <tr> <td>Time in minutes</td> <td>0</td> <td>5</td> <td>10</td> <td>15</td> <td>20</td> <td>25</td> <td>30</td> </tr> <tr> <td>Vol. of gas in cm³</td> <td>0</td> <td>10</td> <td>21</td> <td>26.5</td> <td>31.5</td> <td>32</td> <td>32</td> </tr> </table>	Time in minutes	0	5	10	15	20	25	30	Vol. of gas in cm ³	0	10	21	26.5	31.5	32	32	
Time in minutes	0	5	10	15	20	25	30													
Vol. of gas in cm ³	0	10	21	26.5	31.5	32	32													
	(a)	(i)	What is the role copper (II) sulphate solution?																	
		(ii)	Write ionic equation for the reaction above																	
		(iii)	Explain what would happen to the reaction if zinc granules were replaced with zinc powder.																	
	(b)	(i)	Plot the graph of volume of hydrogen evolved (vertical) against time (horizontal axis)																	
		(ii)	Describe how you would determine the rate of reaction at 12 minutes.																	
		(iii)	Compare the rate of reaction at 12 minutes with that at 20 minutes. Give reason for your answer																	
		(iv)	What happens to the shape of the after 24 minutes? Explain your answer																	

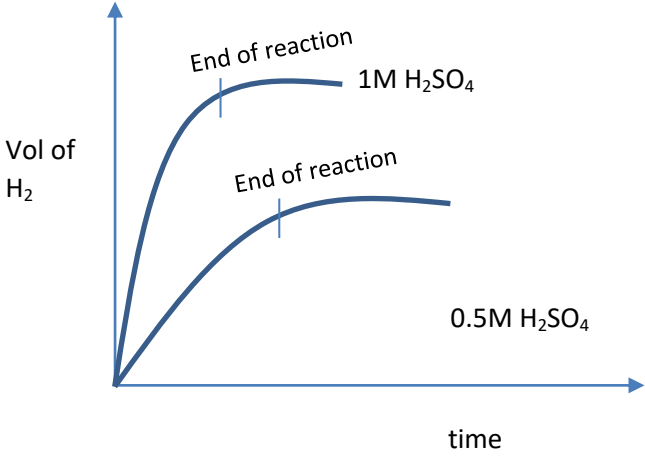
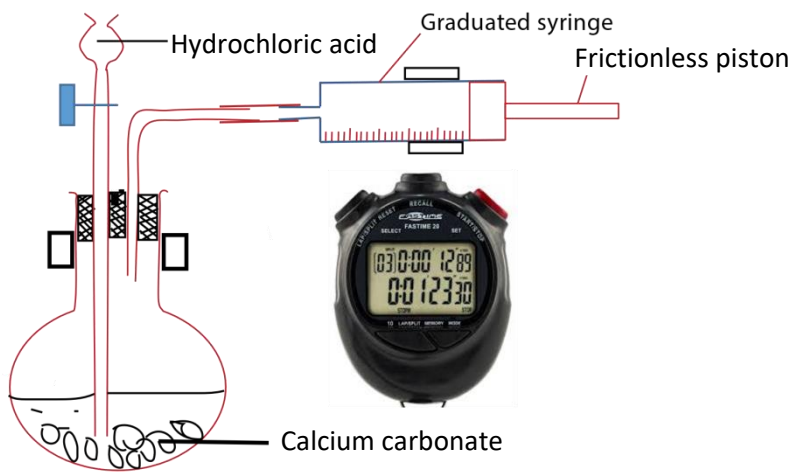
Suggested answers

1.	(a)	(i)	<p>Note that - the rate of reaction in curve X is higher than in curve Y,</p> <ul style="list-style-type: none"> - The initial and final mass in X and Y is the same implying that the mass of iron was not changed <p>Thus, to increase the rate of reaction</p> <ul style="list-style-type: none"> - temperature was increased - particle size of iron was reduced - the concentration acid was increased 						
	(b)	(i)	The concentration of the acid was reduce						
		(ii)	The initial mass is the same but final mass in Z is higher suggesting that not all the iron reacted due to insufficient acid in the volume added.						
2.	(a)		Yellow solid and chocking gas						
	(b)	(i)	Temperature						
		(ii)	The temperature increases the rate of reaction by increasing the rate of collision high energetic molecules.						
		(iii)	<p>Two experiments are carried out using equal volumes of thiosulphate and acid of the same concentration.</p> <p>In one experiment the temperature of thiosulphate is kept at room temperature while for the other the temperature is raised to 60°C.</p> <p>When acid is added to thiosulphate, the yellow suspension form earlier in thiosulphate solution at 60°C that that at room temperature. This indicates that the reaction is faster at higher temperature than at low temperature.</p>						
3.			Refer to page 5-9						
4.		(i)	Time (s)	200	100	40	20	10	
			Concentration of thiosulphate (mol dm ⁻³)	0.05	0.09	0.15	0.20	0.25	
			$\frac{1}{\text{concentration}}$ of thiosulphate (mol ⁻¹ m ⁻³)	20	11	7	5	4	

		(ii)	
		(ii)	<p>Rate of reaction increases as concentration increase</p> <p>(Note that the recipricol of concentration is dilution, and dilute solution takes long to form cloudness)</p>
5.	(a)	(i)	$\text{CaCO}_3(\text{s}) + 2\text{HNO}_3(\text{aq}) \longrightarrow \text{Ca}(\text{NO}_3)_2(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
		(ii)	Temperature increases the rate of reaction
		(iii)	Increase in surface area increases the rate of reaction by increasin the probability of contact between the liquid and the solid
	(b)	(i)	Heat Concentration of hydrogen
	(c)	(i)	
		(ii)	The rate of production of hydrogen increases with the lengnt of magnesium ribblon up to 5cm when the acid become insufficient that increase in rate decreases.
		(iii)	At 4.0 cm the rate is 7.3cm ³ /minute

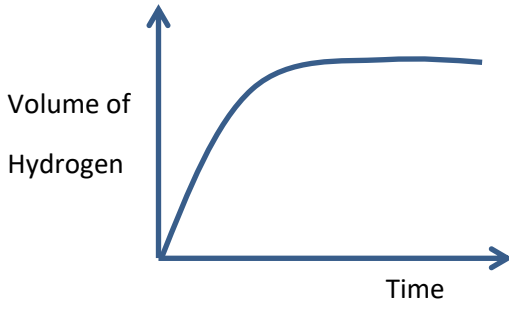
6.	(a)	<p>The setup is as shown below</p>  <p>Hydrogen peroxide is added to manganese (IV) oxide and immediately stop clock started. The amount of oxygen produced in graduated syringe in a given time is noted.</p>
	(ii)	<p>Rate of production of oxygen = $\frac{\text{volume of oxygen}}{\text{time}}$</p>
(b)		 <p>Note that dilution of hydrogen peroxide reduces the rate of production of oxygen and the maximum volume of oxygen produced.</p>
(c)		<p>Mole of hydrogen peroxide = $\frac{200 \times 0.5}{1000} = 0.1 \text{ moles}$</p> <p>Moles of hydrogen produced</p> <p>2mole of hydrogen peroxide produce 1 mole of oxygen</p> <p>$\Rightarrow 0.1 \text{ mole of hydrogen peroxide produce } \frac{0.1 \times 1}{2} = 0.05 \text{ moles}$</p> <p>Mole of magnesium that reacted</p> <p>$2\text{Mg(s)} + \text{O}_2 \text{ (g)} \longrightarrow 2\text{MgO(s)}$</p> <p>From equation</p>

		2 mole of O ₂ react with (2x24 = 48g of magnesium $\Rightarrow 0.05\text{mole of O}_2 \text{ react with } \frac{48 \times 0.05}{2} = 1.2\text{g}$
7	(a)	Rate of reaction is the amount of products produced per unit time
	(b)	
	(d) (i)	At 20 second $\text{Rate} = \frac{47-26}{30-13} = 1.3 \text{ cm}^3/\text{s}$ At 60 seconds $\text{Rate} = \frac{78-66}{70-50} = 0.6 \text{ cm}^3/\text{s}$
	(ii)	The rate at 20 seconds is higher than at 60 seconds because the concentration of the reactants decrease as time goes on.
8	(a) (i)	1000cm ³ contain 0.5 moles $\Rightarrow 100 \text{ cm}^3 \text{ contains } \frac{0.5 \times 100}{1000} = 0.05\text{moles}$
	(ii)	1000cm ³ contain 1 moles $100 \text{ cm}^3 \text{ contains } \frac{1 \times 100}{1000} = 0.1\text{moles}$

	(i) & (ii)	
	(i)	<p>The reaction with 0.5M sulphuric acid takes longer because the reaction is slower. Since the concentration of acid is lower.</p>
9		<p>Experiment set up is shown below</p>  <p>Two experiments are carried out using equal masses of calcium carbonate chips and calcium carbonate powder and equal volumes of acid.</p> <p>In each case the volume of carbon dioxide produced at intervals of time is measured.</p> <p>A plot of the graph of the volume of carbon dioxide by calcium carbonate chips and calcium carbonate powder with time is given below</p>

			<p>Conclusion</p> <p>The graphs show that the rate of reaction is higher with calcium carbonate powder than with chips.</p>
	(b)		The rate of reaction is higher with 4M hydrochloric acid than 2M hydrochloric acid because high concentration of the acid increases the rate of collision between calcium carbonate molecules and acid molecules.
	(c)		Catalyst
10	(a)		The rate of reaction is how fast the reactants are converted to the product with time
	(b)	(i)	<p>Temperature</p> <p>Increasing the temperature increases reaction rates because of the disproportionately large increase in the number of high energy collisions. It is only these collisions (possessing at least the activation energy for the reaction) which result in a reaction.</p>
		(ii)	<p>Surface area:</p> <p>Increasing the surface area of a solid reactant exposes more of its particles to attack. This results in an increased chance of collisions between reactant particles, so there are more collisions in any given time and the rate of reaction increases. The surface of a solid can be increased by crushing it into a powder.</p>

	(c)	(i)	
		(ii)	Rate at 3.0 s = $\frac{77-60}{5.1-2.7} = \frac{17}{2.4} = 7\text{ cm}^3$
	(d)		Would decrease
11	(a)		<p>Experiment set up is shown below</p> <p>Hydrochloric acid is added to iron fillings and immediately a stop clock is started and the volume (V) of hydrogen is recorded in a given time (t)</p> $\text{Rate} = \frac{V (\text{cm}^3)}{t (\text{s})}$
	(b)		$\text{Fe}(\text{s}) + 2\text{HCl}(\text{aq}) \longrightarrow \text{FeCl}_2(\text{aq}) + \text{H}_2(\text{g})$
	(c)	(i)	<p>Moles of HCL that reacted</p> <p>24000cm³ at room temperature of hydrogen require 2 moles of hydrogen</p> <p>120cm³ of hydrogen will require $\frac{2 \times 120}{24000} = 0.01\text{cm}^3$</p> <p>Volume of hydrochloric acid</p> <p>Remember that molarity is the moles in 1000 cm³</p>

		<p>⇒ 0.1moles of HCl are contained in 1000cm^3 And 0.01mole of HCl will be in $\frac{1000 \times 0.01}{0.1} = 100\text{cm}^3$</p>
	(ii)	<p>Mass of iron fillings 2moles of HCl react with 56g ⇒ 0.01 mole of HCl react with $\frac{56 \times 0.01}{2} = 0.28\text{g}$</p>
	(d)	<p>A graph of volume of hydrogen liberated against time</p> 
	(e)	The rate increases

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Thanks

Dr. Bbosa Science