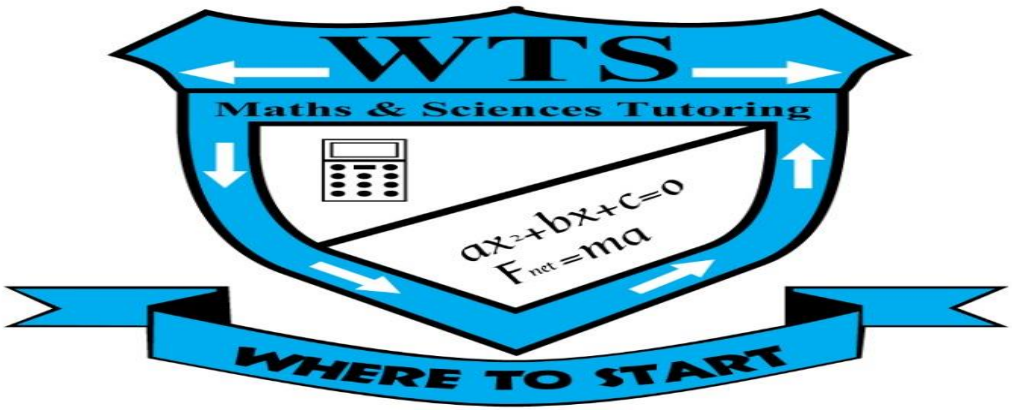
WTS TUTORING



WTS

CHEMICAL CHANGE

GRADE : 12

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PAST PAPERS

QUESTION 5 (Start on a new page.)

Learners investigate the factors that influence the rate of reaction.

In **experiment 1**: One mole of zinc granules reacts with EXCESS dilute hydrochloric acid at 25 °C in an open flask.

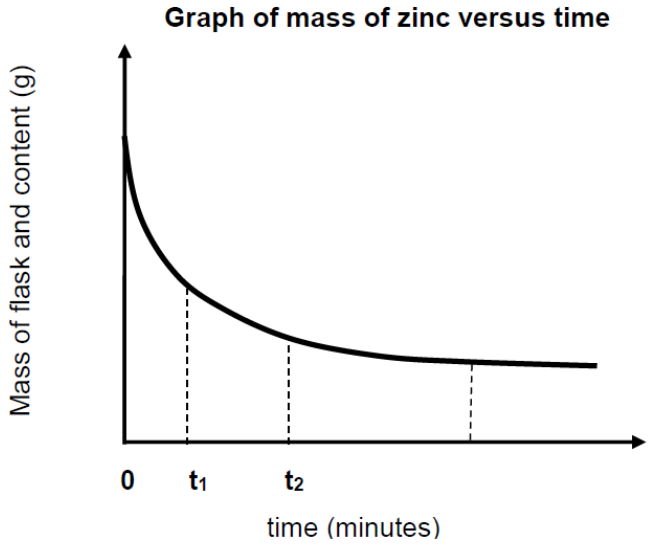
$$Zn\,(s) + 2HC\ell(aq) \longrightarrow ZnC\ell_2(s) + H_2(g) \qquad \Delta H < 0$$

5.1 What change can be made to each of the following in order to increase the rate of production of hydrogen gas?

5.1.1 Zinc granules (1)

5.1.2 Temperature (1)

5.2 A sketch graph of the change in mass of the contents of the flask versus time is shown below.



5.2.1 What does the horizontal part after t_3 indicate about the reaction? (1)

5.2.2 At what time (t_1, t_2 or t_3) is the reaction rate the HIGHEST? Give a reason by referring to the graph. (2)

5.3 Two more experiments (**experiments 2 and 3**) are carried out by changing one condition in **experiment 1**.

In **experiment 2** the zinc granules in **experiment 1** are replaced with one mole of magnesium ribbon.

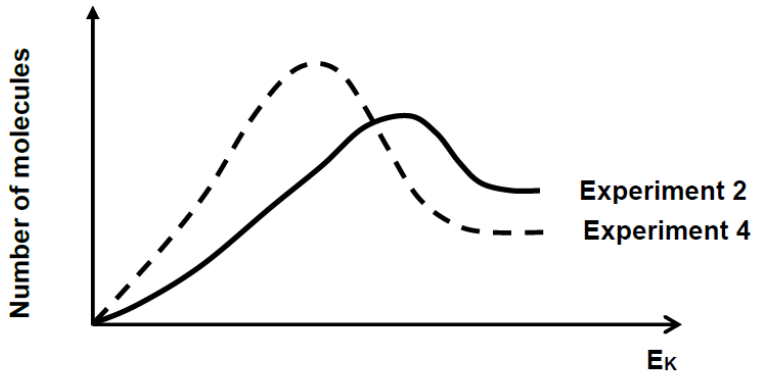
In **experiment 3** zinc granules in **experiment 1** are replaced with one mole of copper turnings.

5.3.1 Redraw the graph in QUESTION 5.2 and label it Zn.
On the same set of axes, sketch a graph that would be obtained for **experiment 2** and label it Mg. (2)

5.3.2 Refer to the relative strengths of oxidising agents from the table of reduction potentials to explain why copper will not react with hydrochloric acid. (2)

5.4 A fourth experiment (**experiment 4**) is carried out by changing one reaction condition in **experiment 2**.

Two energy distribution curves for the reaction in **experiments 2** and **4** are shown in the graph below.



5.4.1 What change was made to the reaction conditions in **experiment 2** to obtain the results of **experiment 4**? (1)

5.4.2 Use the collision theory to explain how the change mentioned in QUESTION 5.4.1 above affects the rate of the reaction. (4)
[14]

QUESTION 6 (Start on a new page.)

The following reversible reaction takes place in a closed container:

$$3\mathbf{X_2}\,(g) + 2\mathbf{A}(s) \rightleftharpoons 2\mathbf{AX_3}\,(g)$$

6.1 Define the term *reversible reaction*. (2)

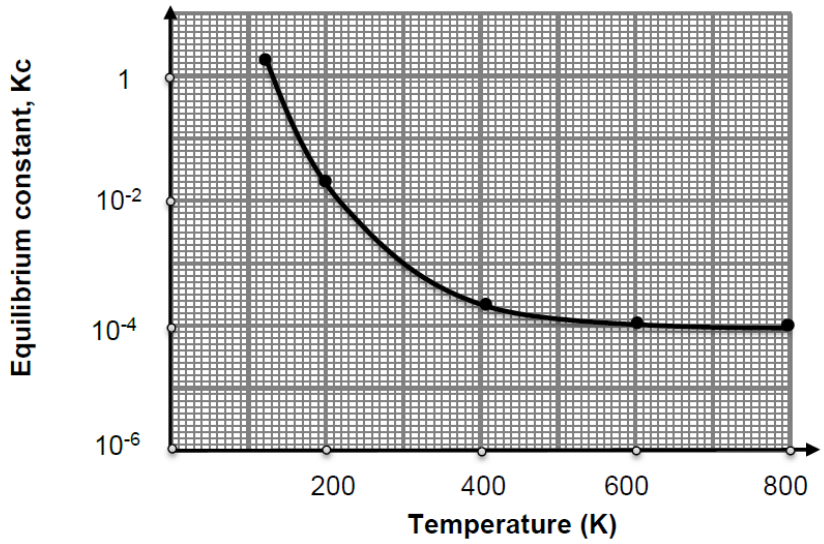
6.2 The reaction reaches equilibrium at a temperature of 300 K.
How will EACH of the following changes affect the amount of $\mathbf{X_2}$?
Choose from INCREASES, DECREASES or REMAINS CONSTANT.

6.2.1 Removing $\mathbf{AX_3}$ as it forms. (1)

6.2.2 Decreasing pressure by increasing the volume. (1)

The graph below shows how the equilibrium constant K_c changes with temperature for the reaction:

Graph of equilibrium constant versus temperature



6.3 Refer to the graph to write down the value of K_c at 300 K. (1)

6.4 When 0,46 moles of $\mathbf{X_2}$ and excess **A** are initially sealed in a container the system reaches equilibrium at 300K.
The concentration of $\mathbf{AX_3}$ at equilibrium is found to be 0,1 mol.dm^{-3}.

Calculate the volume of the container. (8)

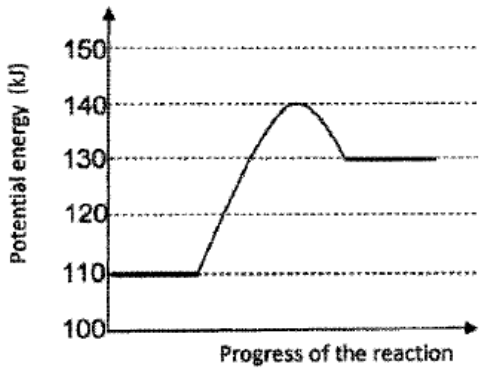
6.5 What effect will an *increase* in temperature have on the amount of $\mathbf{AX_3}$ at equilibrium? ONLY write down INCREASES, DECREASES or REMAINS CONSTANT.

Use Le Chatelier's principle and information from the graph to explain the answer. (4)

[17]

QUESTION 4

4.1 Consider the following potential energy diagram for a reversible reaction.

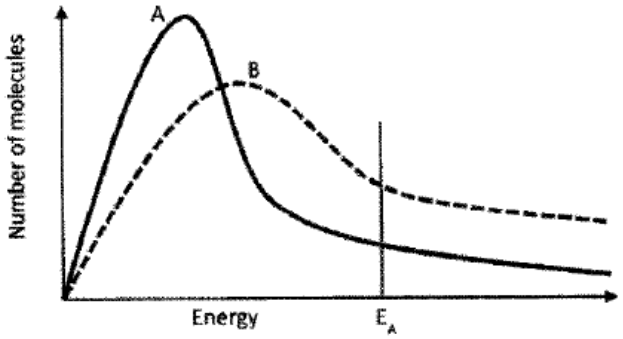


For the **reverse reaction**, write down the value of the ...

4.1.1 *activation energy.* (1)

4.1.2 *heat of reaction.* (1)

4.2 Below is the Maxwell-Boltzman distribution curve of the distribution of the kinetic energy of molecules at two different temperatures.



4.2.1 What does the area under both graphs, **A** and **B**, to the right of the line labelled E_A represent? (1)

4.2.2 One of the reactions takes place at a high temperature. Which ONE of the graphs, **A** or **B**, represents the high temperature? (1)

4.2.3 Explain the answer to Question 4.2.2 in terms of the collision theory. (4)

4.3 The data in the table below indicates the changes in volume of N_2O_5 recorded at different time intervals.

Time (s)	0	100	200	300	400	500	600	700	800
Volume N_2O_5 (cm^3)	0,100	0,081	0,066	0,054	0,044	0,035	0,029	0,023	0,019

4.3.1 Use the table above to draw a graph of the results. Use the attached graph on the ANSWER SHEET at the end of the question paper. (5)

4.3.2 Define the term *reaction rate*. (2)

4.3.3 Calculate the rate of the reaction at t = 240 s. (3)
[18]

QUESTION 5

An equilibrium reaction for the decomposition of a reddish-brown substance, AO_2, is given below. Both products are colourless.

$$4AO_{2(g)} \rightleftharpoons 2A_2O_{3(g)} + O_{2(g)}$$

Brown Colourless

Initially 2,0 mol of A_2O_3 and 1,0 mol of O_2 are present in 1,0 dm^3 container. Only 10,0 % of the AO_2 decomposes at equilibrium.

5.1 Define the term *dynamic equilibrium.* (2)

5.2 Use the information above to calculate the equilibrium concentration of *each* species. (6)

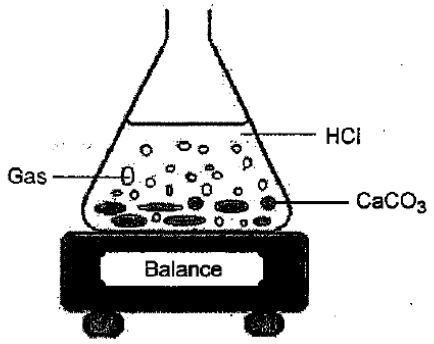
5.3 The volume of the container in Question 5.2 is now reduced to 0,5 dm^3, while the temperature remains constant.

5.3.1 What colour change is observed? Write only BROWN or COLOURLESS. (1)

5.3.2 Use Le Chatelier's Principle to explain the observation made in Question 5.3.1. (3)
[12]

QUESTION 6 (Start on a new page.)

In investigating the factors that affect the rate of reaction, some learners react calcium carbonate with excess dilute hydrochloric acid, in a conical flask. The apparatus is setup as shown below.



The change in mass of the conical flask and its contents are recorded every minute. The results for this investigation are shown in the table below.

Time (s)	0	1	2	3	4	5	6	7	8	9
Mass of flask and its contents (g)	178	176,2	174,1	172,7	172,3	172,2	172,1	172,1	172,1	172,1

6.1 Write down an investigative question for this investigation. (2)

6.2 State the independent variable for this investigation. (1)

The above table is re-drawn in the special answer sheet provided.

6.3 Complete the table by filling in the values for the mass of CO_2 produced. (2)

6.4 Plot a graph of mass of CO_2 produced vs time elapsed. (5)

6.5 Give a reason why the mass of the flask and its contents remains constant after 6 s. (1)

6.6 Calculate the total volume of CO_2 (g) produced at room temperature (25 °C). Assume the molar gas volume of CO_2 (g) at this temperature is 24,46 dm^3. (3)

The experiment is now repeated at a higher temperature. The volume and concentration of the HCl and the mass of $CaCO_3$ used are the same as in the original investigation.

6.7 On the same system of axes used to draw the graph in QUESTION 6.4, draw a sketch graph of mass of CO_2 produced vs time elapsed, at the higher temperature. Label this graph as **N**. (2)

6.8 Use the collision theory to explain the difference between the two graphs obtained. (3)

[19]

QUESTION 7 (Start on a new page.)

Ten (10) grams of hydrogen gas and 355 g of chlorine gas are heated together in a sealed 500 cm^3 container. Equilibrium is reached at 450 °C.

$$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g) \quad ; \quad \Delta H < 0$$

The equilibrium constant for this reaction at 450 °C is 60.

7.1 Calculate the mass of chlorine gas present at equilibrium. (10)

The temperature is now increased to 550 °C while the volume is kept constant. The system reaches a NEW equilibrium.

7.2 State Le Chatelier's principle. (2)

7.3 How will the following be affected in this new equilibrium? Write down only INCREASE, DECREASE or REMAINS THE SAME.

7.3.1 The equilibrium constant. (1)

7.3.2 The volume of H_2 present. (1)

7.4 Use Le Chatelier's principle to explain the answer to QUESTION 7.3.2. (2)

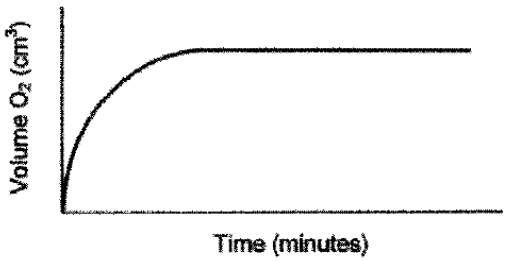
[16]

QUESTION 5 (Start on a new page)

Manganese dioxide (MnO_2) acts as a CATALYST in the decomposition of hydrogen peroxide to produce water and oxygen. The equation for the reaction is as follows:

$$2H_2O_2(aq) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

One gram (1 g) of MnO_2 powder is added to a hydrogen peroxide solution with a concentration of 2 mol·dm^{-3} at 25 °C and the volume $O_2(g)$ which is formed is measured with a gas syringe. The sketch graph below shows the results which were obtained.



5.1 Define the term *catalyst*. (2)

5.2 Redraw the graph in your answer book.

5.2.1 On THE SAME system of axes, sketch the graph that would be obtained if no catalyst is used. Label this graph **A**. (1)

5.2.2 On THE SAME system of axes, sketch the graph that would be obtained if more hydrogen peroxide, at a higher temperature, also with 1 g MnO_2 powder as catalyst, decomposes. Label this graph **B**. (2)

5.3 What mass of MnO_2 will be left at the end of the experiment? (1)

5.4 Define the term *reaction rate*. (2)

5.5 Use the collision theory and explain how a catalyst increases the rate of a reaction. (2)
[10]

QUESTION 6 (Start on a new page)

6.1 Study the reaction below:

$$\underset{\text{(colourless)}}{N_2O_4(g)} \rightleftharpoons \underset{\text{(brown)}}{2NO_2(g)} \quad \Delta H = +58\ kJ{\cdot}mol^{-1}$$

The initial potential energy of the reactants is 20 $kJ{\cdot}mol^{-1}$ and the activation energy for the forward reaction is 90 $kJ{\cdot}mol^{-1}$.

6.1.1 Draw a potential energy diagram for the course of this reaction. Indicate the following on your diagram:
- Energy of the reactants
- Energy of activated complex
- Activation energy of forward reaction (3)

6.1.2 On the same diagram, use a dotted line to indicate the effect of a catalyst on this reaction. (1)

6.1.3 Which reaction will be favoured if the temperature of the reaction mixture is increased? (Write only FORWARD or REVERSE.) (1)

6.1.4 Use Le Chatelier's Principle to explain the answer to QUESTION 6.1.3. (3)

6.2 The hypothetical reaction shown below takes place in a sealed container.

$$A_2(g) + B_2(g) \rightleftharpoons 2AB(g)$$

The change in the equilibrium constant, K_c, at different temperatures, is shown in the table below.

Temperature (°C)	K_c
1000	5,0
750	2,0
500	0,5

Use the information in the table and indicate whether the REVERSE REACTION is exothermic or endothermic. Briefly explain your answer. (3)

6.3 Study the reversible reaction represented by the balanced equation below:

$$H_2(g) + CO_2(g) \rightleftharpoons H_2O(g) + CO(g)$$

Initially x moles $H_2(g)$ are mixed with 0,3 moles $CO_2(g)$ in a sealed 10 dm^3 container. When equilibrium is reached at a certain temperature, it is found that 0,2 moles $H_2O(g)$ are present.

The equilibrium constant (K_c) for this reaction at this temperature is 4.

6.3.1 Calculate the initial number of moles, x, of $H_2(g)$ which were in the container. (9)

6.3.2 The concentration of CO_2 gas is now increased. How does this change affect the value of the equilibrium constant (Kc)? Write only INCREASES, DECREASES or REMAINS THE SAME. (1)

[21]

OUESTION 5 (Start on a new page)

5.1 Define the term *reaction rate.* (2)

5.2 A student investigates the rate of reaction between magnesium and sulphuric acid. The balanced equation for the reaction taking place is:

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

The results are shown in the table below:

Concentration of sulphuric acid ($mol \cdot dm^{-3}$)	Rate of reaction ($cm^3 \cdot s^{-1}$)
1,6	17,0
0,8	8,5
0,4	4,2

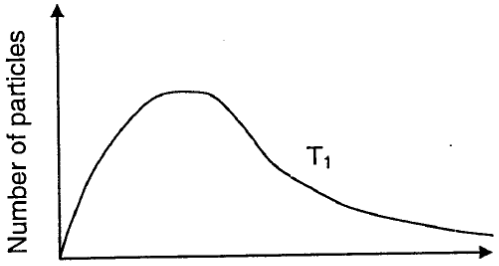
5.2.1 State ONE variable that must be controlled during this investigation. (1)

5.2.2 What conclusion can be drawn from the results obtained. (2)

5.2.3 What will happen to the rate of the reaction if lumps of magnesium are used instead of magnesium powder? Choose from INCREASE, DECREASE or REMAIN THE SAME. (1)

5.2.4 Explain the answer for QUESTION 5.2.3 using the collision theory. (3)

5.2.5 The following diagram shows the distribution curve of the reaction at 20°C.



Redraw the graph in your answer book. Label this graph T_1. Draw a second graph to show how a higher temperature will affect the shape of this graph. Label it T_2. (2)

5.3 Calculate the mass of magnesium needed to produce 100 cm^3 of $H_2(g)$ at 20 °C. The molar gas volume at 20 °C is 24,04 cm^3. (5)
[16]

QUESTION 6 (Start on a new page)

The reaction below represents the catalysed step in the Haber process.

$$N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g) \qquad \Delta H<0$$

The reaction takes place in a closed container and reaches equilibrium at 427 °C.

6.1 What will the influence of a drop in temperature be on the following:
Choose from INCREASES, DECREASES or REMAIN THE SAME.

6.1.1 The rate of the forward reaction. (1)

6.1.2 The yield of $NH_3(g)$. (1)

6.1.3 Explain the answer in QUESTION 6.1.2 by using Le Chatelier's principle. (3)

6.2 The reaction is investigated on a small scale in the laboratory. Initially 4 mole $N_2(g)$ and an unknown mass, *x*, of $H_2(g)$ are sealed in a 2 dm^3-flask and allowed to reach equilibrium at a certain temperature.

At equilibrium, the concentration of $NH_3(g)$ present in the flask is 1,5 $mol{\cdot}dm^{-3}$. Calculate the initial mass of $H_2(g)$ present in the flask if the equilibrium constant (K_c) at this temperature is 1,8. (8)

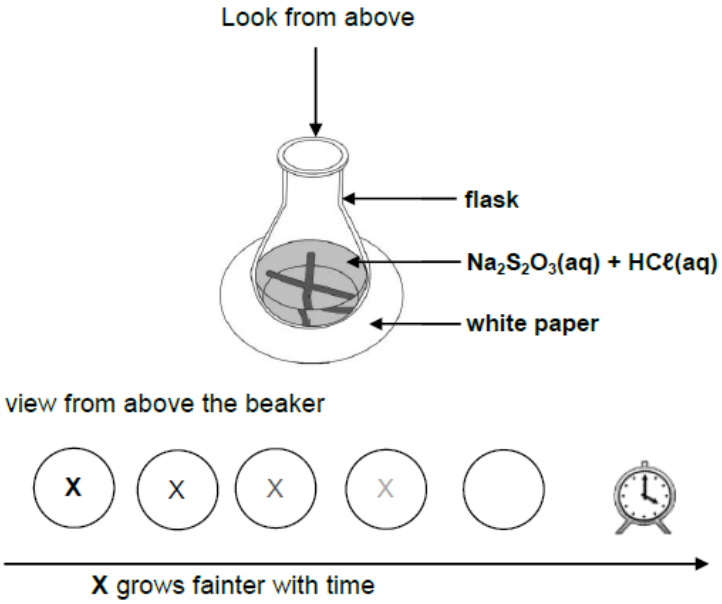
[13]

QUESTION 5 (Start on a new page.)

A group of learners were investigating the factors affecting the reaction rate. In one of their reactions sodium thiosulphate ($Na_2S_2O_3$) and hydrochloric acid ($HC\ell$) are reacted according to the balanced equation given below:

$$Na_2S_2O_3(aq) + 2HC\ell(aq) \rightarrow 2NaC\ell(aq) + S(s) + H_2O(\ell) + SO_2(g)$$

During the reaction one of the products formed causes the solution to turn cloudy. Due to this, the mark **X** drawn on paper, on which the reaction mixture is placed, were not visible after the reaction is completed. Refer the sketch below.



In the investigation the learners took sodium thiosulphate of different concentration and reacted it with hydrochloric acid of constant concentration. They recorded the time taken for the mark **X** on the paper to disappear. Their result is tabulated below.

Trial	1	2	3	4
Concentration $Na_2S_2O_3$(mol.dm^{-3})	0,15	0,09	0,06	0,03
Time taken for the X to disappear in (s)	43	66	100	240

5.1 Define the term *reaction rate*. (2)

5.2 For the experiment described above, name the:

5.2.1 Independent variable (1)

5.2.2 ONE control variable (1)

5.3 Write down the NAME or FORMULA of the substance responsible for the cloudiness. (1)

5.4 Consider the table of results and write down which trial (1, 2, 3 or 4) occurred at the highest reaction rate? (1)

5.5 Draw a graph of concentration of sodium thiosulphate against time for the mark **X** to disappear in the ATTACHED GRAPH SHEET. (4)

5.6 Use the graph to make a conclusion for the above investigation. (2)

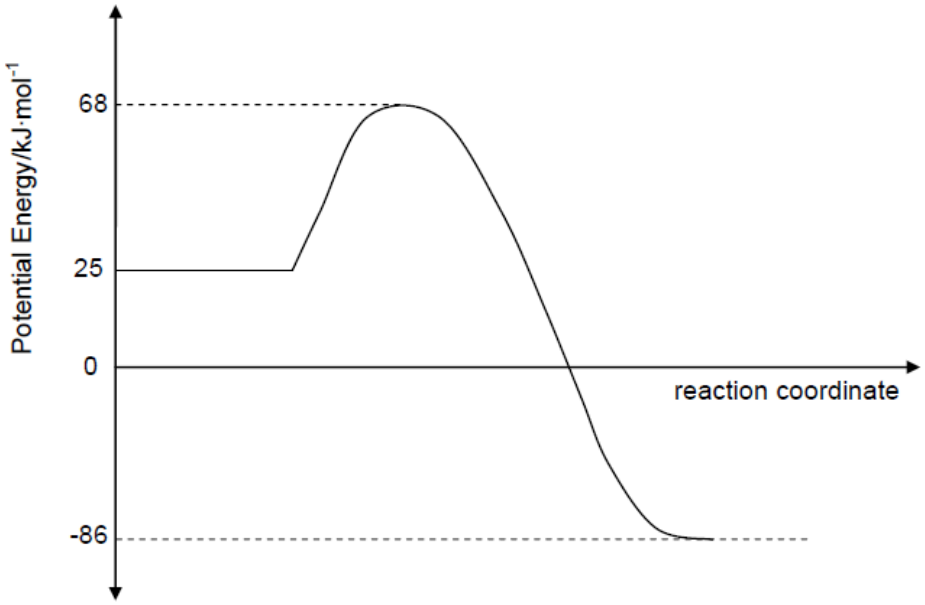
5.7 Name TWO other ways in which the rate of this reaction can be increased. (2)
[14]

QUESTION 6 (Start on a new page.)

One of the steps in the preparation of sulphuric acid in the industry is represented by the following reversible reaction:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

The graph below shows the energy change during this reaction.



6.1 Write down the type of reaction represented by above graph. Choose from EXOTHERMIC or ENDOTHERMIC. Explain your answer. (2)

6.2 According to collision theory, give TWO conditions necessary for the reaction to take place. (2)

Vanadium pentoxide is added as a catalyst in the above reaction.

6.3 Explain how the presence of a vanadium pentoxide as a catalyst affect the rate of above reaction. (2)

6.4 Calculate the enthalpy change of this reaction. (3)

At 68 $kJ \cdot mol^{-1}$ an activated complex is formed.

6.5 Define the term *activated complex.* (2)

6.6 Calculate the activation energy for the reverse reaction. (2)

[13]

QUESTION 7 (Start on a new page.)

Consider the reversible reaction represented by the following balanced equation.

$$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$$

During this reaction 4 mole of nitrogen monoxide gas, 2,5 mole of oxygen and ***x*** mole of nitrogen dioxide gas is placed in an 500 cm^3 container at a temperature of 450 K. At equilibrium there is 3 mole of nitrogen oxide present in the container. The equilibrium constant (K_c) at this temperature is 0,25.

7.1 Calculate the initial mole of nitrogen dioxide placed in the container. (7)

7.2 Is the given equilibrium system an example of a heterogeneous or homogeneous equilibrium? Explain your answer. (2)

7.3 State Le Chatelier's principle. (2)

Now the pressure is increased.

7.4 Use Le Chatelier's principle to explain how this change affects the concentration of NO_2 at the new equilibrium. (2)

[13]

QUESTION 5 (Start on a new page.)

An excess diluted $HC\ell(aq)$ is used to dissolve 5 g of $CaCO_3(s)$ in a beaker during an experiment on reaction rate.The following reaction takes place:

$$CaCO_3(s) + 2HC\ell(aq) \rightarrow CaC\ell_2(aq) + H_2O(\ell) + CO_2(g)$$

The results obtained are summarised in the table below:

Mass of beaker, $HC\ell$ and $CaCO_3(s)$	62,0	61,2	60,6	60,2	60,0	60,0
Time (s)	0	10	20	30	40	50

5.1 Give a reason why the mass of the beaker with its contents decreased in the first 30 s. (1)

5.2 How long did it take to use up all the $CaCO_3(s)$ in this experiment? Explain how you arrived at the answer. (2)

5.3 Calculate the volume of $CO_2(g)$ produced at STP. (6)

5.4 When the diluted $HC\ell(aq)$ is replaced by concentrated $HC\ell(aq)$, the rate of the reaction increases. Use the collision theory to explain this observation. (3)
[12]

QUESTION 6 (Start on a new page.)

The Haber process is represented by the following equation:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H = -92\ kJ$$

In a small scale plant, **x** moles of $N_2(g)$ and 8 moles of $H_2(g)$ are added in a 5 dm^3 sealed container. When equilibrium is reached at temperature T_1, it is found that 25,5 g of $NH_3(g)$ is present.

6.1 Define the term *chemical equilibrium.* (2)

6.2 If the K_c value for this reaction is 0,13, calculate the initial number of moles of $N_2(g)$. (9)

6.3 How will the equilibrium concentration of the product compare with that of the reactants? Choose from LARGER THAN, SMALLER THAN or EQUAL TO. Refer to the provided Kc value in QUESTION 6.2 to give a reason for the answer. (2)

6.4 A new equilibrium is now established at a higher temperature T_2. Will the value of the equilibrium constant (K_c) INCREASE, DECREASE or STAY THE SAME?
Use Le Chatelier's principle to explain the answer. (4)
[17]

QUESTION 5 (Start on a NEW PAGE.)

Learners investigate some of the factors that influence the rate of a chemical reaction. In the experiment they add equal amounts of each of three different metals separately to equal volumes of EXCESS dilute hydrochloric acid solution.

In each experiment the acid completely covers the metal.

The data obtained is recorded as in the table below:

Experiment	Amount of metal powder	Change in temperature of solution (°C) ($T_{final} - T_{initial}$)	Time taken to run to completion
1	0,1 mol Zn	+23	25,2
2	0,1 mol Mg	+37	8,3
3	0,1 mol Cu	0	No reaction

5.1 Is the reaction in **Experiment 1** ENDOTHERMIC or EXOTHERMIC? Give a reason for your answer. (Use the information in the table). (2)

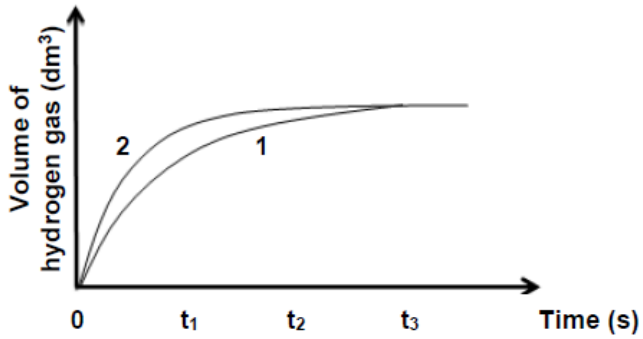
5.2 Which factor influencing reaction rate is investigated? (1)

5.3 How will the total volume of hydrogen gas produced in **Experiment 2** compare with the total volume of hydrogen gas produced in **Experiment 1** at the end of the reactions?

Write down HIGHER THAN, EQUAL TO or SMALLER THAN.

Give a reason for your answer. (2)

5.4 The graphs obtained for **Experiment 1** and **Experiment 2**, labelled as **1** and **2** respectively, are sketched on the same set of axes as shown below:



5.4.1 In which experiment does the reaction occur at a higher reaction rate at time t_1? (1)

5.4.2 Explain the answer to QUESTION 5.4.1 by referring to the relative strength of reducing agents involved. (2)

5.5 In another experiment, **Experiment 4**, the same reaction conditions are repeated as in **Experiment 2,** but the reaction mixture is heated. The rate of reaction is HIGHER for **Experiment 4** than **2**.

Explain why the reaction rate is HIGHER for **Experiment 4** than **2** by referring to the collision theory. (3)

[11]

QUESTION 6 (Start on a NEW PAGE.)

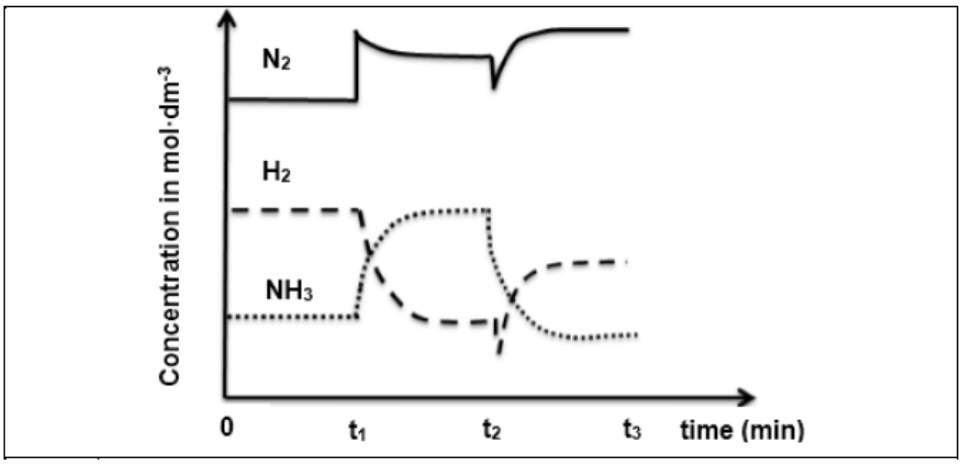
The reaction represented by the equation below reaches equilibrium in a closed container.

$$N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g) \qquad \Delta H = -92 \text{ kJ}$$

6.1 Is the above equilibrium HOMOGENEOUS or HETEROGENEOUS?

Give a reason for the answer. (2)

6.2 Changes were made to the TEMPERATURE, PRESSURE and CONCENTRATION of the above equilibrium mixture. The graphs below represent the results obtained.



6.2.1 What changes were made to the reaction conditions at each of the following times?

(a) t_1 (1)

(b) t_2 (1)

6.2.2 How does the rate of the forward reaction compare to the rate of the reverse reaction between **0** and t_1?

Write down HIGHER THAN, LOWER THAN or EQUAL TO. (1)

6.3 Equal number of moles of hydrogen gas and nitrogen gas are injected into a sealed 1 dm^3 container. When the reaction reaches equilibrium at temperature T_1 it is found that 10% of the original amount of hydrogen is ***left*** in the container. The value of K_c at temperature T_1 is $1{,}426 \times 10^3$.

6.3.1 Calculate the initial mass of N_2 in the container. (10)

6.3.2 Use your knowledge of Le Chatelier's principle to explain how an increase in temperature will affect the value of K_c. (3)

[18]

QUESTION 5 (Start on a new page.)

In an investigation of the rate of reaction, excess magnesium powder is added to dilute hydrochloric acid at room temperature. The following spontaneous reaction takes place:

$Mg(s) + 2HC\ell(aq) \rightarrow MgC\ell_2(aq) + H_2(g)$ $\Delta H < 0$

5.1 Define the term *spontaneous reaction*. (2)

5.2 Write down the limiting reagent for the above reaction. (1)

5.3 How will each of the following changes affect the rate of the reaction between magnesium and hydrochloric acid according to the above reaction? Choose from INCREASES, DECREASE or REMAINS THE SAME.

5.3.1 The same mass of magnesium ribbon is used instead of powder. (1)

5.3.2 A more concentrated solution of hydrochloric acid is used. (1)

5.3.3 The diluted hydrochloric acid solution is heated before being added to the magnesium. (1)

5.4 Use the collision theory to explain your answer in QUESTION 5.3.3. (3)
[9]

QUESTION 6 (Start on a new page.)

During the industrial preparation of ammonia, nitrogen gas and hydrogen gas react in a closed container until the following equilibrium is established at a constant temperature of 472 °C. Kc is 0,1 at this temperature of 472 °C. The volume of the container is 0,5 dm^3.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H = -92\ kJ{\cdot}mol^{-1}$$

The equilibrium concentrations are:

$[NH_3] = 2,7 \times 10^{-3}\ mol{\cdot}dm^{-3}$
$[H_2] = 1,221 \times 10^{-1}\ mol{\cdot}dm^{-3}$

6.1 Write down the term for the underlined phrase. (1)

6.2 Write down the NAME or FORMULA of the catalyst used in this reaction. (1)

6.3 After equilibrium has been established, the temperature remained constant. Explain this observation. (2)

6.4 Calculate the initial mass of nitrogen gas. (10)

6.5 Explain why such a high temperature is used although the yield is low. (3)

[17]

QUESTION 5

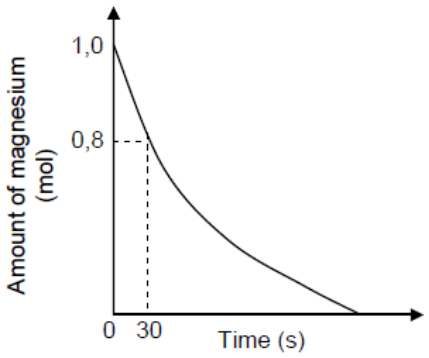
A group of learners use the reaction between magnesium and nitric acid to investigate one of the factors that affects reaction rate. The reaction that takes place is represented by the balanced equation below.

$$Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$$

They add magnesium ribbon to *dilute* nitric acid and measure the mass of magnesium used per unit time. The experiment is repeated using *concentrated* nitric acid.

5.1 Write down an investigative question for this investigation. (2)

The results obtained for the reaction with dilute nitric acid are represented in the graph below:



5.2 Which substance, **Mg** or **HNO_3**, is in excess? Use the information on the graph to give a reason for the answer. (2)

5.3 Define the term *reaction rate*. (2)

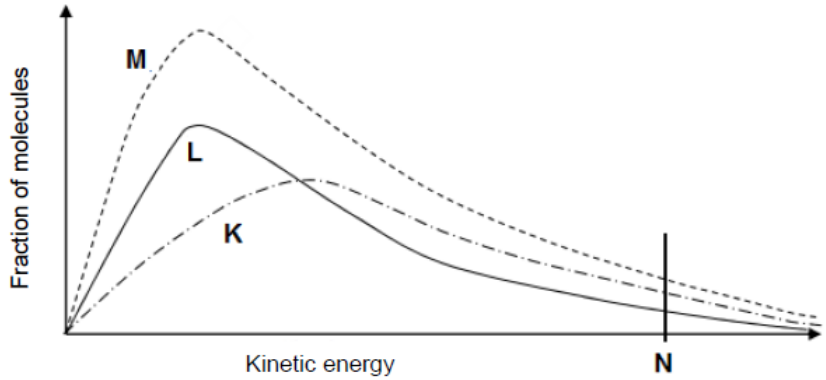
5.4 Calculate the average rate of the reaction (in gram per second) during the first 30 s. (5)

5.5 Copy the above graph in your answer book. On the same set of axes use a **DOTTED LINE** to show the curve that will be obtained when concentrated nitric acid is used. No numerical values are required. (2)

[13]

QUESTION 6

6.1 Curve **L** shown below, is the Maxwell-Boltzman distribution curve for a gas in a closed container at 250 °C.



6.1.1 Name the energy represented by **N**. (1)

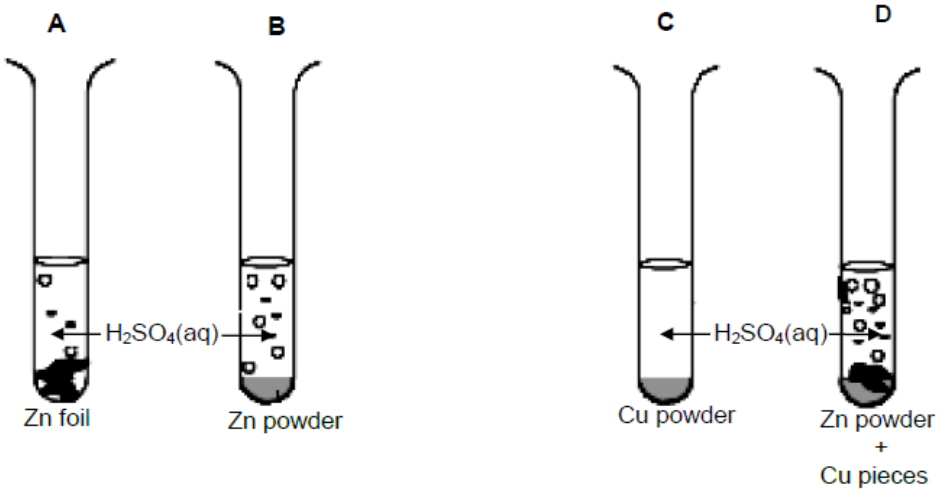
6.1.2 Write down the change in reaction conditions that resulted in:

(a) Curve **M**. (1)

(b) Curve **K**.

Use the collision theory to explain how this change affects the rate of the reaction. (4)

6.2 A series of experiments are carried out to compare the reactions of zinc foil, zinc powder, copper powder and a mixture of zinc powder and copper pieces with dilute sulfuric acid of concentration 1 mol·dm^{-3}. Hydrogen gas is produced in all test tubes where a reaction takes place. The diagram shows the test tubes some time after the metals have been added to the acid.



6.2.1 Refer to the relative strengths of oxidizing agents or reducing agents to explain why no reaction takes place in test tube **C**. (3)

6.2.2 How does the rate of the reaction in test tube **B** compare to that in:
(Choose from GREATER THAN, SMALLER THAN or EQUAL TO.)

(a) Test tube **A**?
Give a reason for your answer.
(2)

(b) Test tube **D**?
Give a reason for your answer. (2)

[13]

QUESTION 7

7.1 The reaction represented below reaches equilibrium in a closed container.

$$4C(s) + 5H_2(g) \rightleftharpoons C_4H_{10}(g)$$

The equilibrium constants for this reaction at two different temperatures are given in the table below.

TEMPERATURE (K)	EQUILIBRIUM CONSTANT (K_c)
400	$1,58 \times 10^{-3}$
600	$1,58 \times 10^{-9}$

7.1.1 Is the forward reaction ENDOTHERMIC or EXOTHERMIC? (1)

7.1.2 Use Le Chatelier's principle to explain your answer in QUESTION 7.1.1. (3)

7.1.3 The pressure in the container is now decreased by increasing the volume of the container. What effect will this have on the value of the equilibrium constant?

Choose from INCREASES, DECREASES or REMAINS THE SAME. (1)

7.1.4 Give a reason for the answer to QUESTION 7.1.3. (1)

7.2 Exactly 24,0 mol $SO_3(g)$ is sealed in an empty 2,0 dm^3 container. The reaction reaches equilibrium at 700 K after 8 minutes according to the following balanced equation.

$$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$$

If the reaction mixture contains 10,0 mol $O_2(g)$ at equilibrium at 700 K, calculate the equilibrium constant (K_c) at 700 K. (7)

[13]

QUESTION 5 (Start on a new page)

Ketiwe conducts an experiment to investigate the various factors that influence the rate of chemical reactions. She places a sample of calcium carbonate in a beaker. The beaker is placed on a sensitive balance and an excess of hydrochloric acid (HCl) is added.

The following reaction occurs:

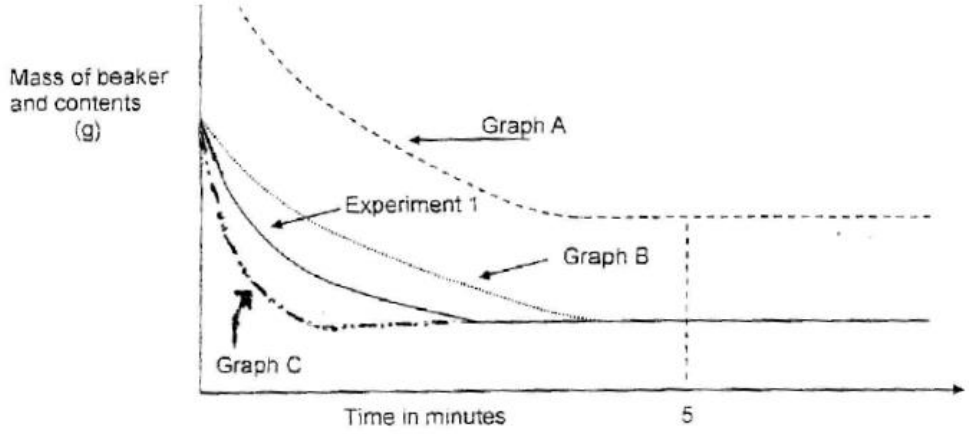
$$CaCO_3 + HCl \rightarrow CaCl_2 + H_2O + CO_2$$

Ketiwe repeats the experiment a number of times under different conditions, always with the same volume of HCl, which remains in excess.

The following table summarizes the different experimental conditions for four of her experiments.

Expt	Mass of $CaCO_3$ (g)	Concentration of HCl ($mol \cdot dm^{-3}$)	Temperature of HCl (^oC)	State of $CaCO_3$
1	10	2	25	granules
2	10	2	15	granules
3	20	2	25	granules
4	10	2	25	powder

During each experiment the mass of the beaker and its contents is recorded every minute. The graphs below indicate the changes in mass of the beaker and its contents during the reaction, as a function of time, for the four experiments:



5.1 What is meant by the rate of a chemical reaction? (2)

5.2 Name the quantity that was kept constant during all 4 experiments. (1)

5.3 Give a reason for the decrease in mass as each reaction progresses. (2)

5.4 Why are all the graphs horizontal lines after five minutes? (2)

5.5 Which one of the graphs (A, B, or C) represents the results of:

5.5.1 Experiment 2 (1)

5.5.2 Experiment 3 (1)

5.5.3 Experiment 4 (1)

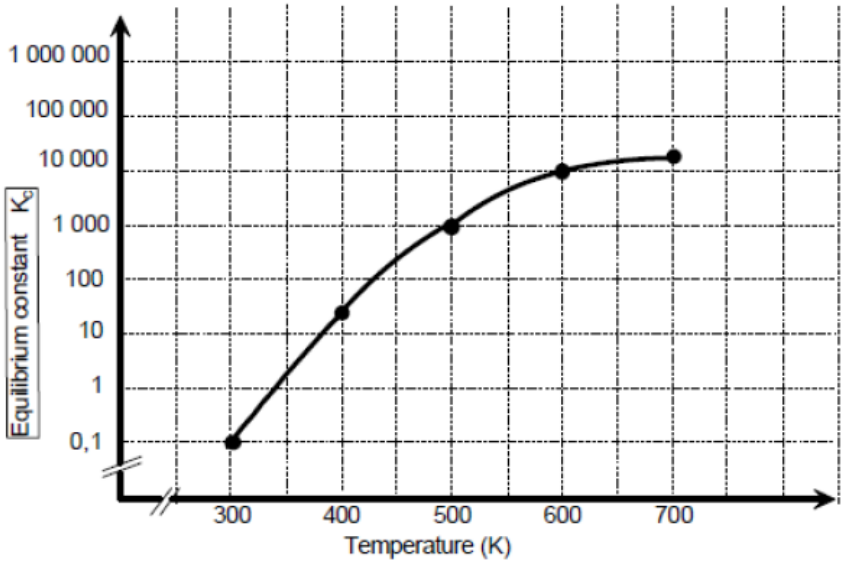
5.6 If a suitable catalyst is used in experiment 1, which of the graphs (A,B or C) will be obtained? Explain your answer. (4)

[14]

QUESTION 6 (Start on a new page)

6.1 The graph below shows the effect of a temperature change on the value of K_c for the following reaction taking place in a closed container:

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$



6.1.1 What effect does an increase in temperature have on the amount of NO_2 formed? (1)

6.1.2 Which reaction was favoured due to an increase in temperature?
Write only FORWARD or REVERSE. (1)

6.1.3 State Le Chatelier's Principle. (2)

6.1.4 Using Le Chatelier's Principle, explain whether the forward reaction is EXOTHERMIC or ENDOTHERMIC. (4)

6.1.5 Write down TWO factors, other than temperature, that can be used to increase the rate of the forward reaction at 500K. (2)

6.2 Consider the hypothetical reaction that takes place between A_2 and B in a closed container.

$$\underset{\text{colourless}}{A_2(g)} + \underset{\text{colourless}}{2B(g)} \rightleftharpoons \underset{\text{dark red}}{2AB(g)} \qquad \Delta H > 0$$

X mol of gas A_2 and 2,0 mol of gas B are sealed in a 1,0 dm^3 container. After a few minutes equilibrium is reached and the contents of the container turns light red.

At equilibrium it is found that 0,40 mol of gas AB is present in the container. The value of K_c is 0,50.
Determine X, the quantity (in mol) of gas A_2 that was originally sealed in the container. (8)

[18]

QUESTION 5

5.1 The amount of product formed/reactant used up✓ per unit time✓ (per second). (2)

5.2 concentration of HCl✓ (1)

5.3 Product CO_2 is an insoluble gas✓ and escapes from the beaker. ✓ (2)

5.4 Calcium carbonate is used up✓ and the reaction stops, no more CO_2 formed. ✓ (2)

5.5.1 graph B✓ (1)

5.5.2 graph D ✓ (1)

5.5.3 graph C✓ (1)

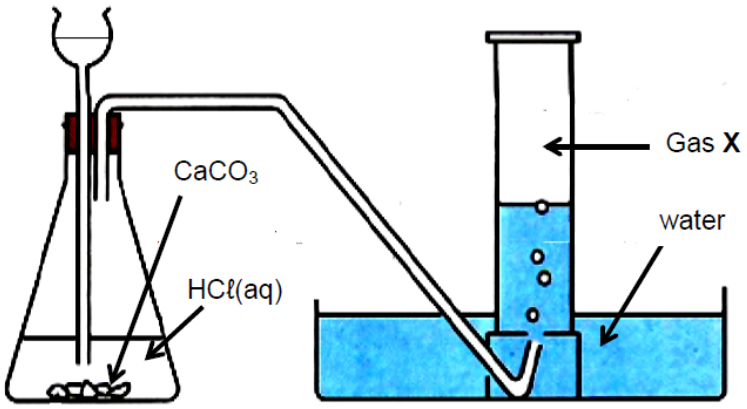
5.6 graph C✓.
Catalyst speeds up the rate of reaction✓, therefore gradient ✓of graph will be steepest for same initial mass. ✓ (4)

5.7 Remains the same. ✓ (1)

[15]

QUESTION 5 (Start on a New Page)

A group of grade 12 learners uses the reaction between calcium carbonate and hydrochloric acid to investigate one of the factors that influence reaction rate. They use the apparatus shown below.



The reaction that takes place is represented by the following chemical equation:

$$CaCO_3(s) + 2HC\ell(aq) \rightarrow CaC\ell_2(aq) + \mathbf{X}(g) + H_2O(\ell) \quad \Delta H < 0$$

5.1 Identify the gas **X.** (2)

5.2 Two experiments are conducted by using the apparatus shown above.

The conditions for each experiment are given in the table below:

Experiment	Mass of $CaCO_3(s)$ (g)	State of division of $CaCO_3(s)$	Concentration of $HC\ell$ ($mol{\cdot}dm^{-3}$)	Temperature of $HC\ell(aq)$ ($^{\circ}C$)
1	4	lumps	0,2	40
2	4	lumps	0,4	40

5.2.1 Define, in words, the term *reaction rate* by referring to the concentration of hydrochloric acid. (2)

5.2.2 FROM THE TABLE ABOVE, write down the independent variable for this investigation. (1)

5.2.3 Give a reason why the learners use equal masses and the same state of division of $CaCO_3(s)$. (1)

The learners observe that the reaction rate is HIGHER in **experiment 2** than in **experiment 1.**

5.2.4 Use the collision theory to explain this observation. (3)

5.2.5 Refer to **experiment 2** and calculate the volume of hydrochloric acid (in cm^3) that reacts with $CaCO_3(s)$. Assume that $CaCO_3$ is the LIMITING REAGENT. (4)

5.3 Sketch a POTENTIAL ENERGY versus REACTION COORDINATE graph for this reaction.

Label the axes and indicate the following on the graph:

- Heat of reaction
- Activation energy
- Activated complex

(4)
[17]

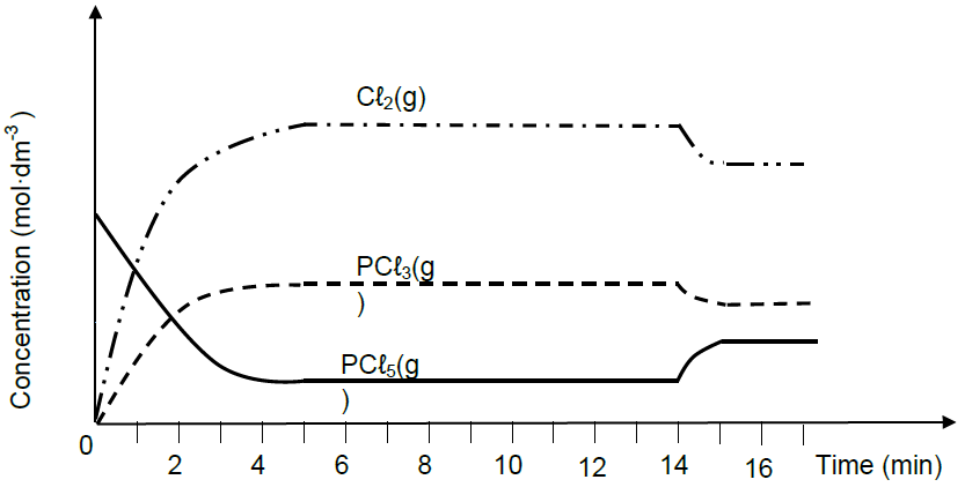
QUESTION 6 (Start on a New Page)

A company uses the decomposition of phosphorus pentachloride, $PC\ell_5(g)$, at 200 °C to produce phosphorous trichloride, $PC\ell_3(g)$.

The reaction which takes place in a sealed container is represented by the balanced equation below:

$$PC\ell_5(g) \rightleftharpoons PC\ell_3(g) + C\ell_2(g)$$

The results of the adjustments are shown in the graph below.



6.1 How much time was required for the system to come to chemical equilibrium for the FIRST time? (1)

6.2 How does the rate of the FORWARD reaction compare to that of the REVERSE reaction during the following time intervals:
(Choose from GREATER THAN, EQUAL TO or LESS THAN)

6.2.1 0,0 – 2,0 minutes (1)
6.2.2 8,0 – 10,0 minutes (1)

6.3 Initially Phosphorus pentachloride, $PC\ell_5(g)$ of unknown mass *x* is injected into an empty 2 dm^3 container at 200°C. At equilibrium, it is found that the concentration of $PC\ell_5(g)$ is 0,15 $mol \cdot dm^{-3}$.
The K_c for this reaction at 200°C is 5,55.

Calculate the initial mass of $PC\ell_5(g)$ gas injected into the container. (9)

6.4 If the change at t = 14 minutes is due to a decrease in temperature state whether the FORWARD reaction is ENDOTHERMIC or EXOTHERMIC.

Use Le Chatelier's principle to explain the answer. (3)
[15]

QUESTION 5 (START ON A NEW PAGE)

A chemist wishes to determine the rate of reaction of zinc with hydrochloric acid. The unbalanced equation for the reaction is:

$$Zn(s) + HC\ell(aq) \rightarrow ZnC\ell_2(aq) + H_2(g)$$

A piece of zinc is dropped into 0,1 cm^3 of 0,1 $mol \cdot dm^{-3}$ $HC\ell$ and the following data were obtained at 4 second intervals.

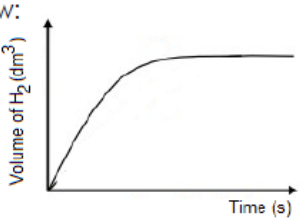
Time (s)	Mass of zinc (g)
0	0,016
4	0,012
8	0,010
12	0,009
16	0,008
20	0,008

5.1 Calculate the average rate of the reaction for the first 12 s in $mol \cdot s^{-1}$. (4)

5.2 Explain why the mass of the zinc remained constant after 16 s. (1)

5.3 Explain how the rate of the reaction changes as the time passes. Answer only INCREASES, DECREASES or REMAINS THE SAME. (1)

5.4 A graph that shows the amount of H_2 gas that was produced against time for this reaction is shown below:

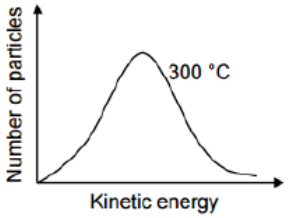


Redraw the graph and indicate on the same set of axes

5.4.1 a second graph, labelled **Y** that will be obtained if zinc powder of the same mass, instead of zinc granules was used. (2)

5.4.2 a third graph, labelled **X** that will be obtained if the same volume of $HC\ell$ with a **lower** concentration was used. (2)

The Maxwell-Boltzmann distribution curve below represents the number of particles against kinetic energy at 300 °C.



5.5 Redraw this curve in the ANSWER BOOK. On the same set of axes, sketch the curve that will be obtained at a temperature of 400 °C. Clearly label the curves as 300 °C and 400 °C respectively. (2)

5.6 Using the collision theory, explain how an increase in temperature affects the rate of reaction. (2)

[14]

QUESTION 6 (START ON A NEW PAGE)

33,6 g of N_2 and 24 g of H_2 are placed in a 5 dm^3 sealed container and react according to the following balanced equation until dynamic equilibrium is reached.

$$N_2(g) + 3\ H_2(g) \rightleftharpoons 2NH_3(g)$$

6.1 Explain the term "*dynamic chemical equilibrium*" (2)

At 573 K when equilibrium is reached, there is 5,6 g of N_2 gas left inside the container.

6.2 Calculate the concentration of both reactants when the reaction reached equilibrium for the first time. (6)

6.3 Calculate the equilibrium constant (K_c) when this reaction reached equilibrium. (3)

6.4 How would the yield of NH_3 change if a smaller container than the 5 dm^3 is used? Write only INCREASES, DECREASES or REMAINS THE SAME. (1)

The temperature of this reaction increases from 573 K to 700 K. At this new temperature the equilibrium constant for this reaction decreases.

6.5 Is the forward reaction endothermic or exothermic? (1)

6.6 Explain your answer to QUESTION 6.5. (3)

[16]

QUESTION 5 (Start on a new page.)

5.1 Define the term *reaction rate* in words. (2)

Learners use the reaction between IMPURE POWDERED zinc and excess hydrochloric acid to investigate reaction rate. The balanced equation for the reaction is:

$$Zn(s) + 2HC\ell(aq) \rightarrow ZnC\ell_2(aq) + H_2(g)$$

They perform four experiments under different conditions of concentration, mass and temperature as shown in the table below. They use identical apparatus in the four experiments and measure the volume of gas released in each experiment.

	EXPERIMENT			
	1	**2**	**3**	**4**
Concentration of acid ($mol \cdot dm^{-3}$)	1	0,5	1	1
Mass of impure zinc powder (g)	15	15	15	25
Initial temperature of acid (°C)	30	30	40	40

5.2 The results of experiments **1** and **3** are compared in one of these investigation.

Write down the:

5.2.1 Independent variable (1)

5.2.2 dependent variable for this investigation. (1)

5.3 Use the collision theory to explain why the reaction rate in experiment **1** will be higher than that in experiment **2**. (3)

5.4 Experiment 3 and experiment 4 are now compared with each other.

5.4.1 How will the reaction rate of experiment **3** compare to that of experiment **4**? Write down only HIGHER THAN, LOWER THAN or EQUAL TO. (1)

5.4.2 Draw a sketch graph to show the difference between experiment **3** and **4**. Place the volume H_2 gas on the y axis and time on the x axis. No values have to be indicated on the graph. Clearly mark the line graphs as **experiment 3** and **experiment 4**. (4)

5.5 When the reaction in experiment 4 reaches completion, the volume of the gas formed is 8,6 dm^3. Determine the percentage purity of the zinc powder. PLEASE NOTE: The molar gas volume at 40° C is equal to 25,7 dm^3. (NOT 22,4 dm^3) (5)

[17]

QUESTION 6 (Start on a new page.)

Nitrosyl bromide decomposes to form nitrogen(II) oxide and bromide gas according to the balanced equation below:

$$2NOBr\,(g) \rightleftarrows 2NO\,(g) + Br_2\,(g)$$

55g NOBr is sealed in a 2 dm^3 container and allowed to decompose. At equilibrium 78% of the NOBr has decomposed.

6.1 Calculate the K_c – value for this reaction. (8)

6.2 The same reaction takes place at the **same temperature**, but in a 1 dm^3 container. How will the following be influenced?
(Write only INCREASE, DECREASE or STAY THE SAME.)

6.2.1 K_c value. (2)

6.2.2 Time it takes to reach equilibrium. (2)

6.2.3 The number of moles of Br_2 (2)

6.3 Use Le Chatelier's principle to fully explain your answer in 6.2.3. (3)
[17]

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