

ACDB4011: Rate of Reaction

UNIVERSITY OF WALES TRINITY ST DAVID – OME PROGRAMME
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1.0 INTRODUCTION

Purpose

The purpose of the experiment is to react magnesium metal with hydrochloric acid under different conditions and record the volume of gas produced over specific time periods. By doing this experiment, previous knowledge of reactions can be validated showing how temperature and rate of reaction are closely related, and that the surface area of the metal and concentration (molarity) of hydrochloric acid will determine how much Hydrogen gas is produced over certain time intervals.

Hypotheses to be Tested

When magnesium is reacted with hydrochloric acid, hydrogen gas will be produced as well as magnesium chloride which will become a solute within the HCl solution. Based on prior knowledge of the reactivity of magnesium it is likely that the gas being produced will be of a large enough quantity to be captured over a couple of minutes.

The rate of chemical reactions increases with temperature, so it is expected that the temperature of the HCl solution is increased, the rate of reaction will also increase. This hypothesis will be tested in Experiment 1.

The molarity (concentration) of the HCl will affect the speed of the reaction. The higher the molarity (concentration) of HCl, the faster it is expected for the reaction to occur. This hypothesis will be tested in Experiment 2.

Surface area can affect a chemical reaction. Usually, a larger surface area in an experiment would cause the rate of reaction to increase. This hypothesis is tested in Experiment 3.

Equations

The rate of reaction equation for this experiment will be:-

$$\text{rate} = \Delta V_{\text{H}_2} / \Delta t [\text{cm}^3/\text{s}]$$

Were:-

ΔV is change in volume

ΔT is change in time

Variables

Independent variables. The lab experiment was to examine the differences in rate of reaction using different molarity (concentration) of acid and the temperature of the solution. By keeping the mass of magnesium the same and changing the molarity (concentration) of acid or temperature, the volume of hydrogen and the rate of gas produced was collected and analysed.

Dependant variables.

The dependant variable was the volume of Hydrogen gas captured during the reaction of Magnesium and the Hydrochloric acid.

Controlled variables. These were conditions kept constant throughout the experiments. In these experiments the controlled conditions were the ambient temperature and pressure although these variables were not specifically controlled but were assumed to be constant throughout the experiment.

2.0 THEORY

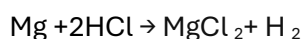
Basic concepts

The Law of the conservation of energy states that matter cannot be destroyed nor created but only changed in state. This is why chemical equations are balanced so that each molecule or atom is accounted for even if it is in a different state of matter [3].

There are three common acids that are used in these types of experiments: Hydrochloric acid (HCl), Nitric acid (HNO₃) and Sulfuric acid (H₂SO₄). When a metal reacts with an acid, a metal salt will be produced and also hydrogen gas.

Chemical equation and formulas for calculations

There is one chemical equation for the formation of Hydrogen gas and magnesium chloride. The balanced equation is:



Examples of applications/importance of kinetics in OME

An understanding of the rate of reaction is critical when applied to OME because the rapid decomposition of materials to produce large quantities of gas is what causes explosions to have a huge amount of destructive power. For example, rockets are fuelled by pyrotechnic compositions which decompose at a subsonic rate [4]. If a material such as RDX was used as a rocket motor, the decomposition of the material would likely become a detonation which means the reaction occurs faster than the speed of sound which would not only result in damage to the projectile and rocket motor but it would not have the decomposition duration required to propel the rocket from A to B.

3.0 RESULTS

Experiment 1 – Effect of HCl concentration on the speed of H₂ gas evolution

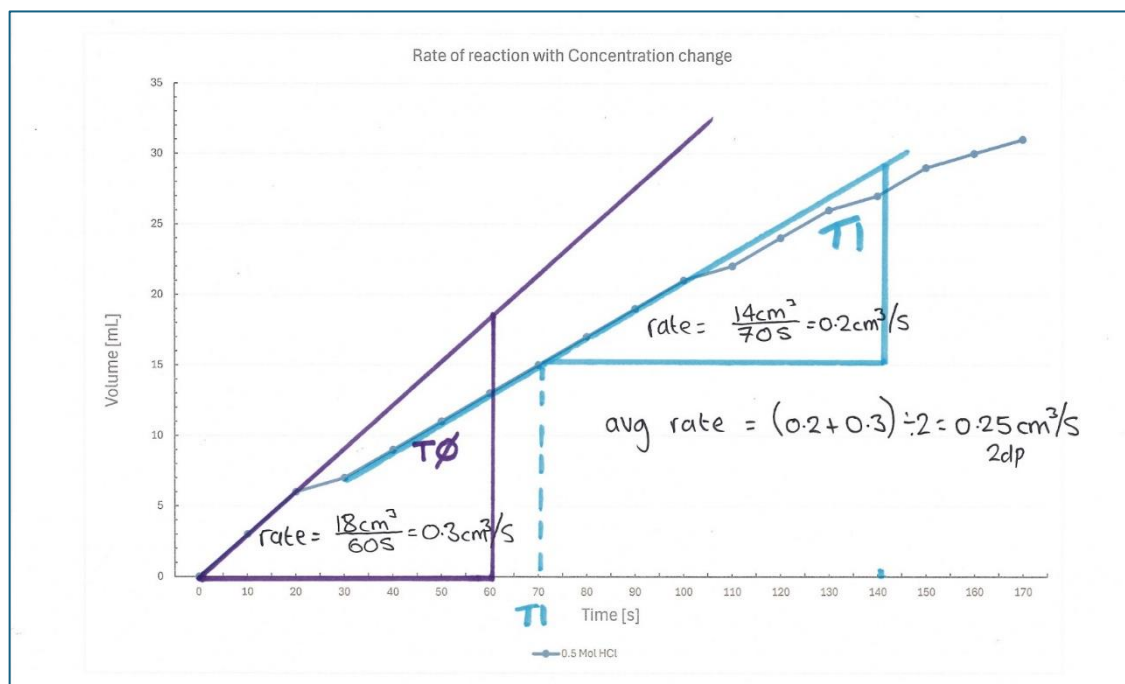


Figure 1 Experiment 1, 0.5 Mol HCl

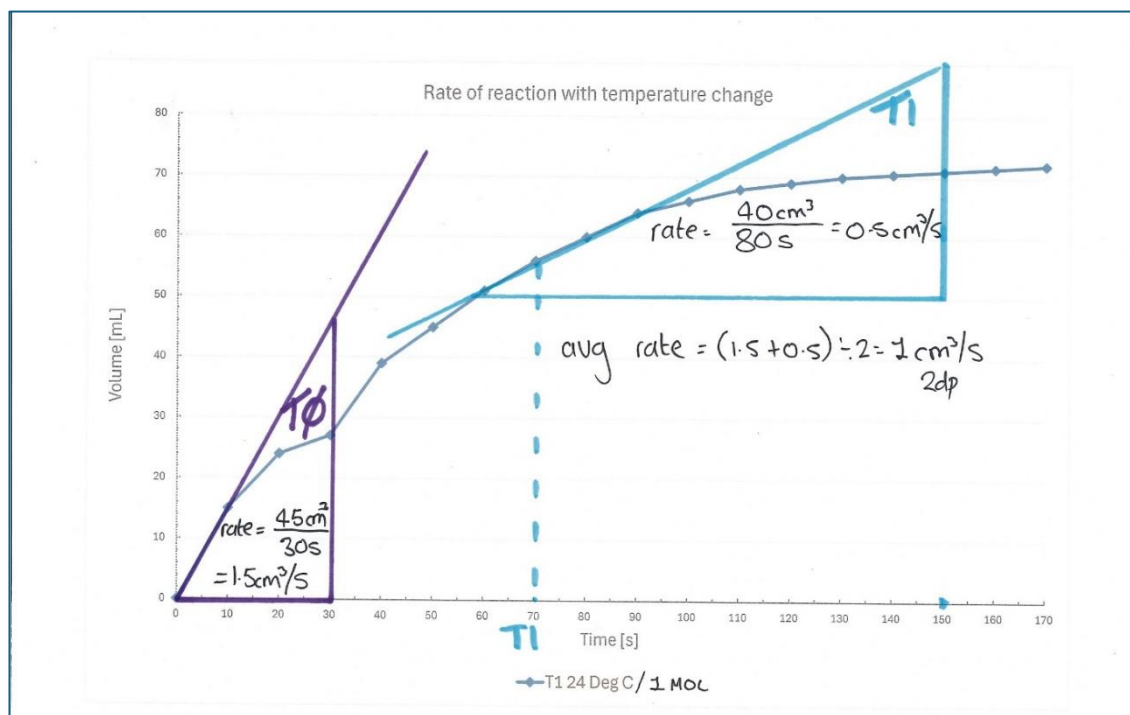


Figure 2 Experiment 1, 1.0 Mol HCl

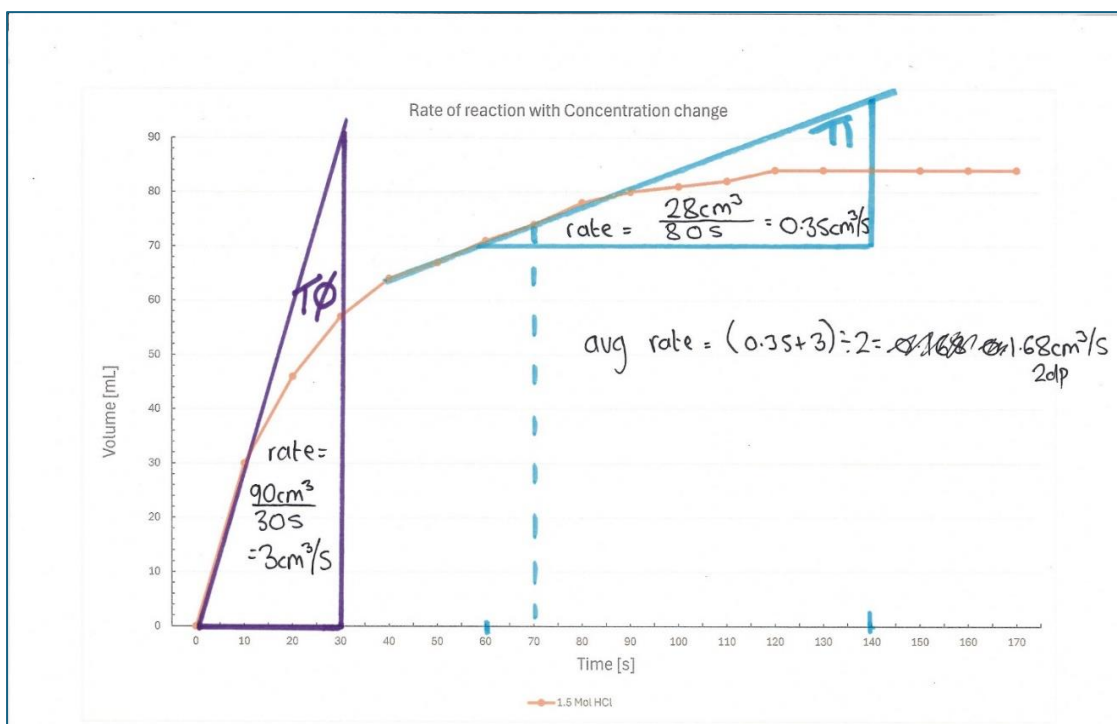


Figure 3 Experiment 1, 1.5 Mol HCl

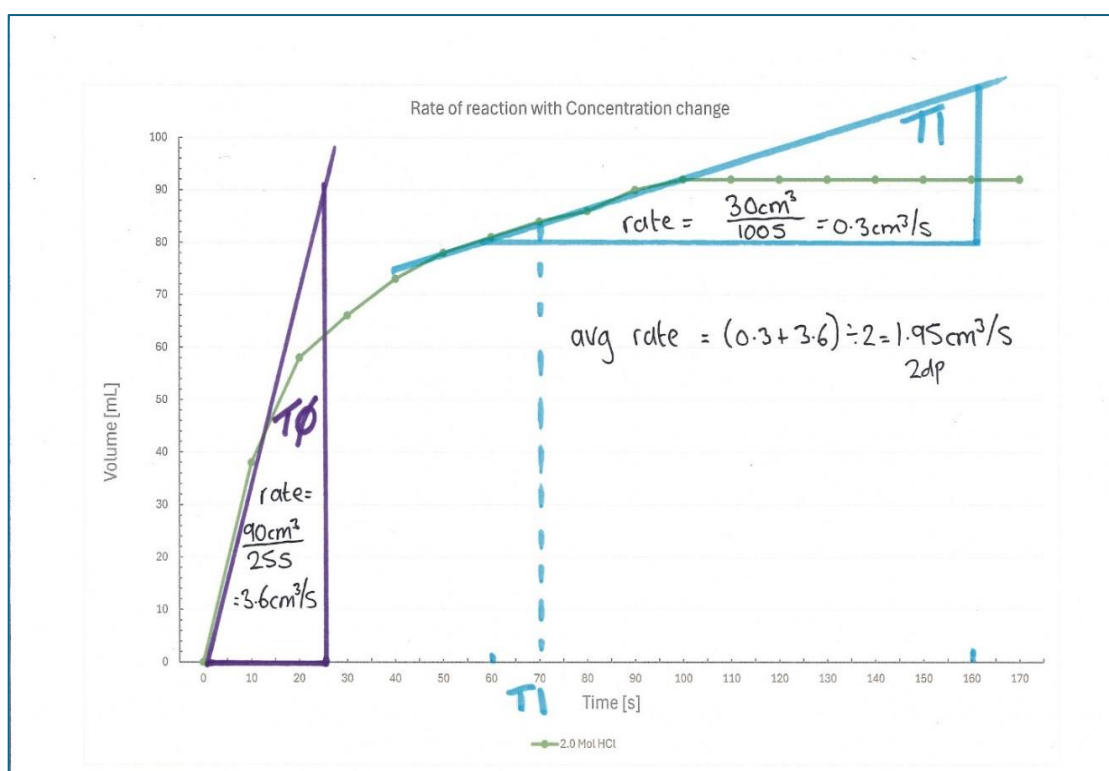


Figure 4 Experiment 1, 2.0 Mol HCl

Experiment 1			
HCl	Time to dissolve Mg [s]	Average Rate [cm^3/s]	Instantaneous Rate @ $t=0$ [cm^3/s]
0.50	170	0.25	0.30
1.00	170	1.00	1.50
1.50	120	1.68	3.00
2.00	100	1.95	3.60

Table 1 – Summary of Experiment 1 Results

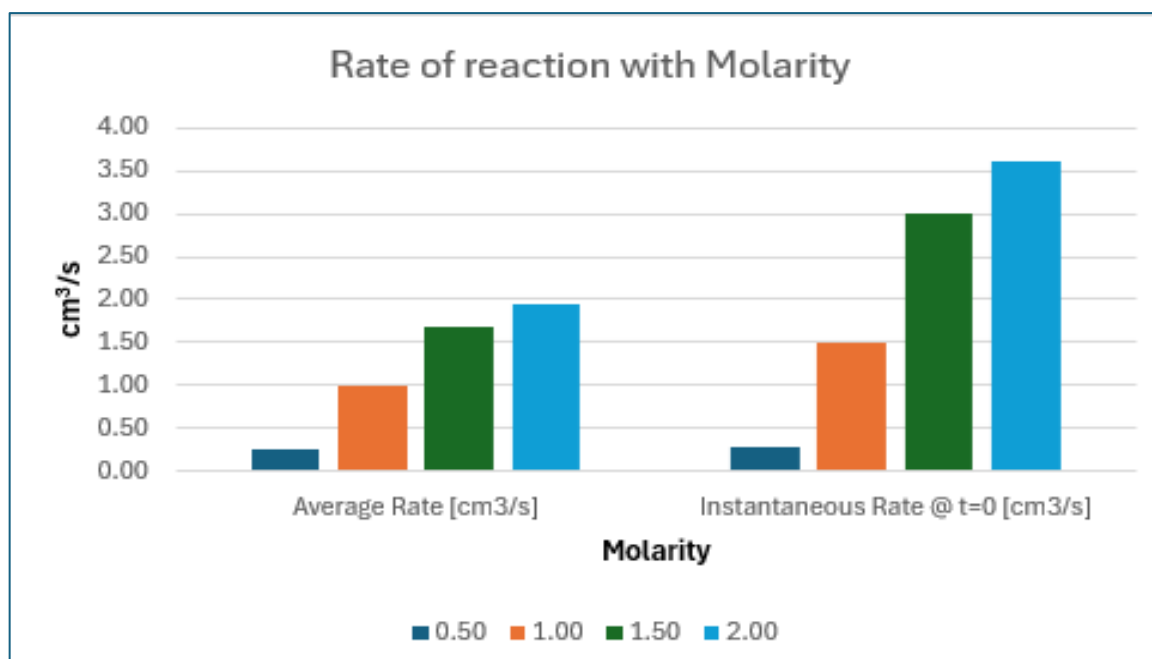


Figure 5 – Bar Chart Showing Rate of Reaction Between Different Molarity (Concentration) of HCl

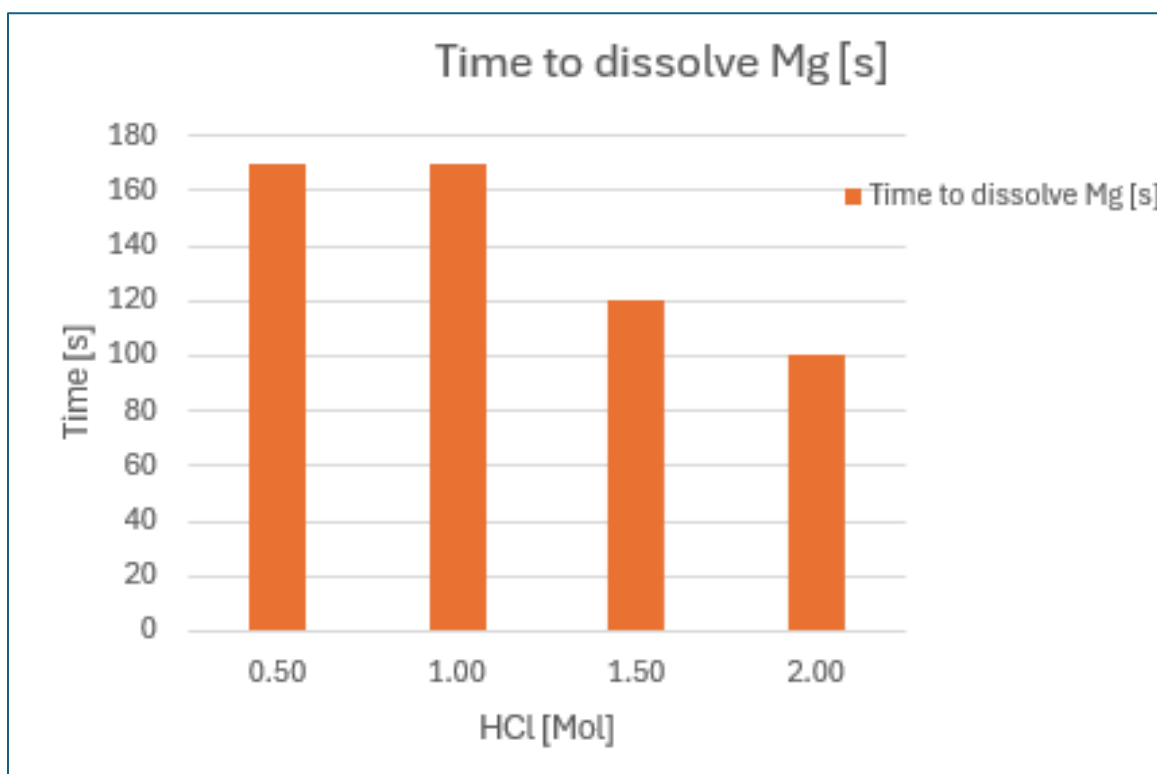


Figure 6 – Bar Chart Showing Time to Fully Dissolve Mg Between Different Molarity (Concentration) of HCl

Experiment 1 Findings

By showing the results using graphs the rate of reaction and time taken to dissolve the magnesium is proportional to the molarity (concentration) of HCl. Table 2 shows that as molarity (concentration) increases, so does the initial rate of reaction and also the average rate.

Table 3 shows that although the overall time taken to fully dissolve the magnesium strips reduces as molarity (concentration) increases, the time differences are not significant especially between 0.5 – 1.0 Mol of HCl.

Experiment 2 – Effect on rate of reaction with temperature change

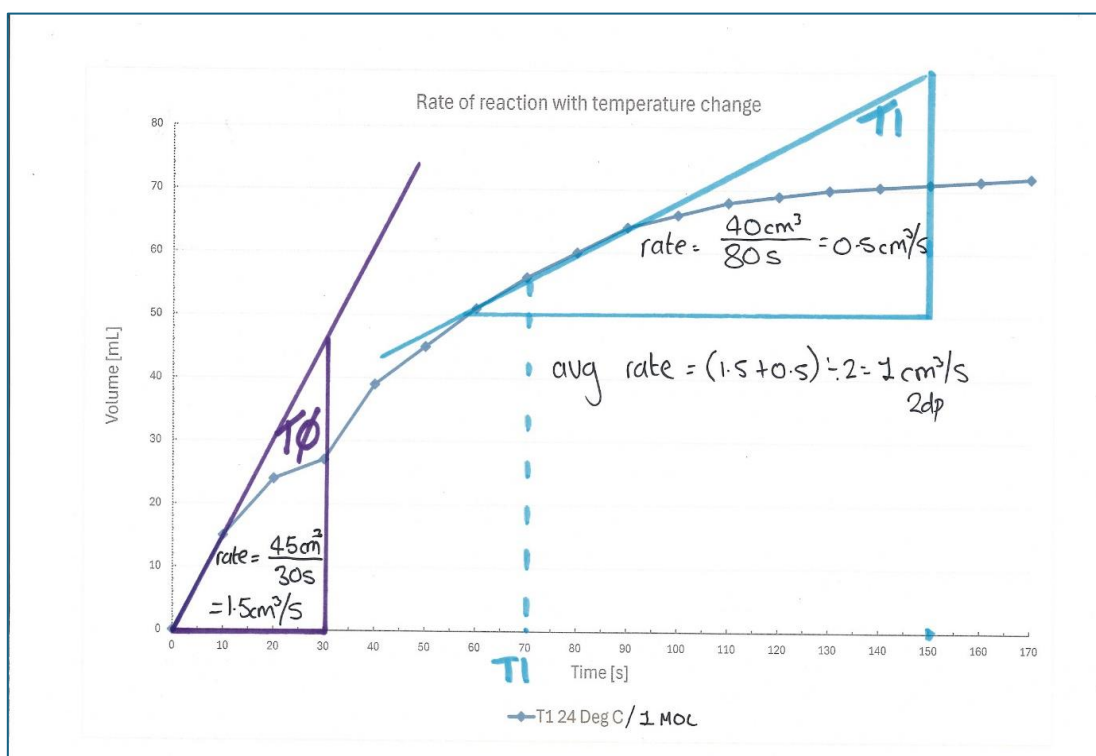


Figure 7 Experiment 2, 1.0 Mol at 24°C

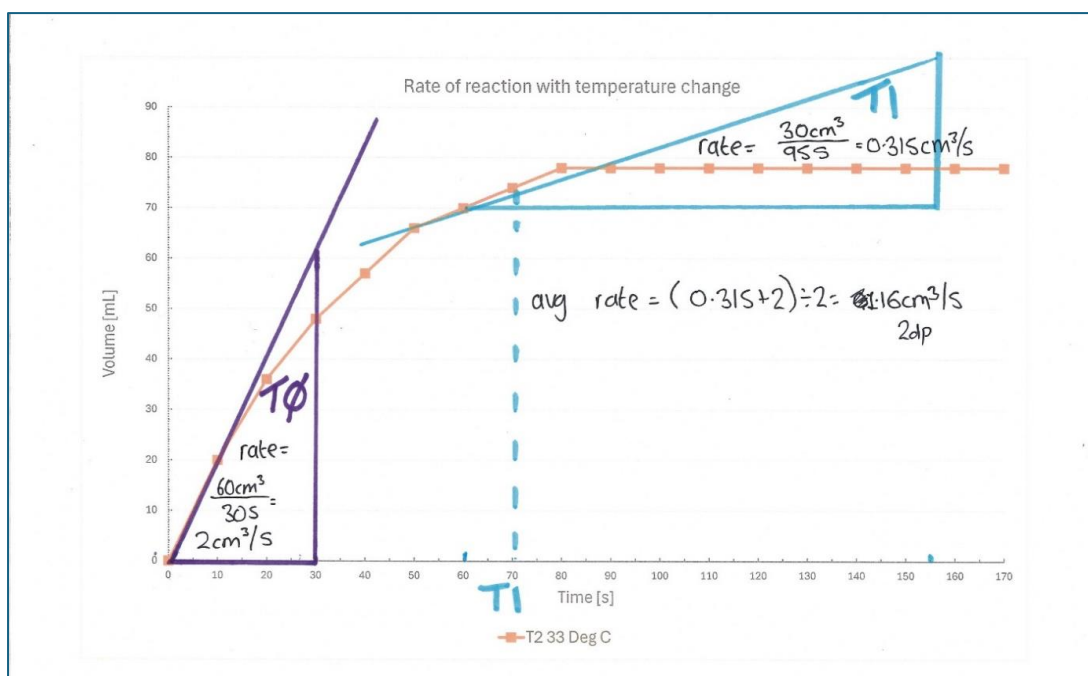


Figure 8 Experiment 2, 1.0 Mol at 33°C

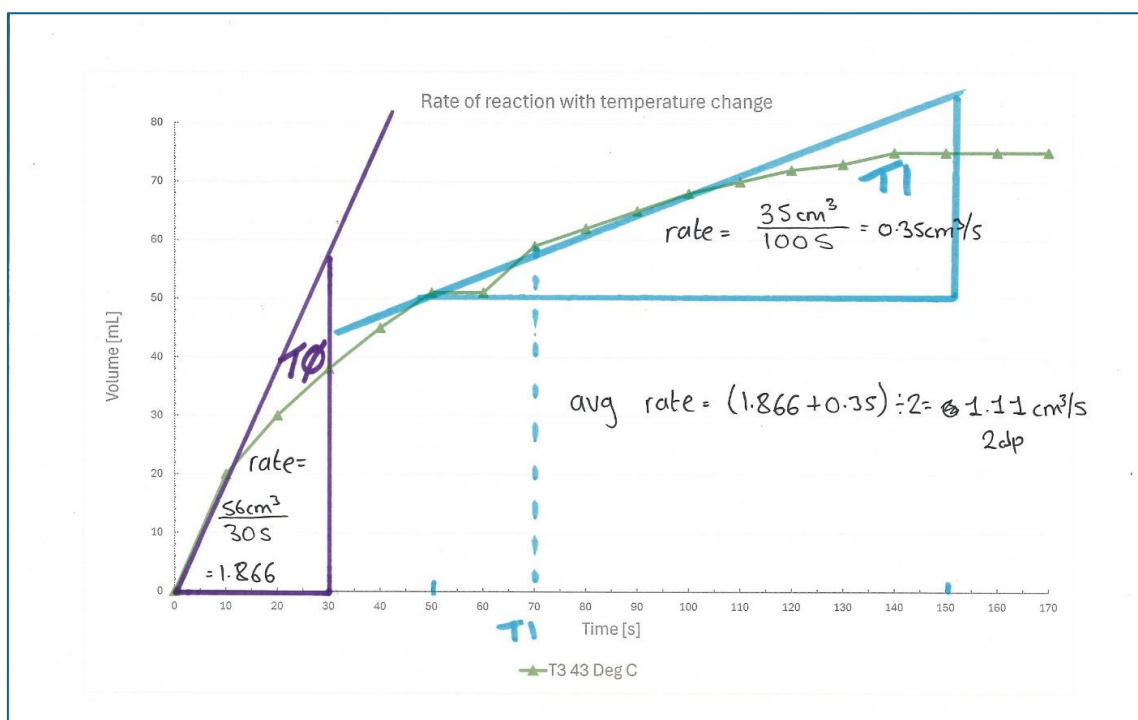


Figure 9 Experiment 2, 1.0 Mol at 43°C

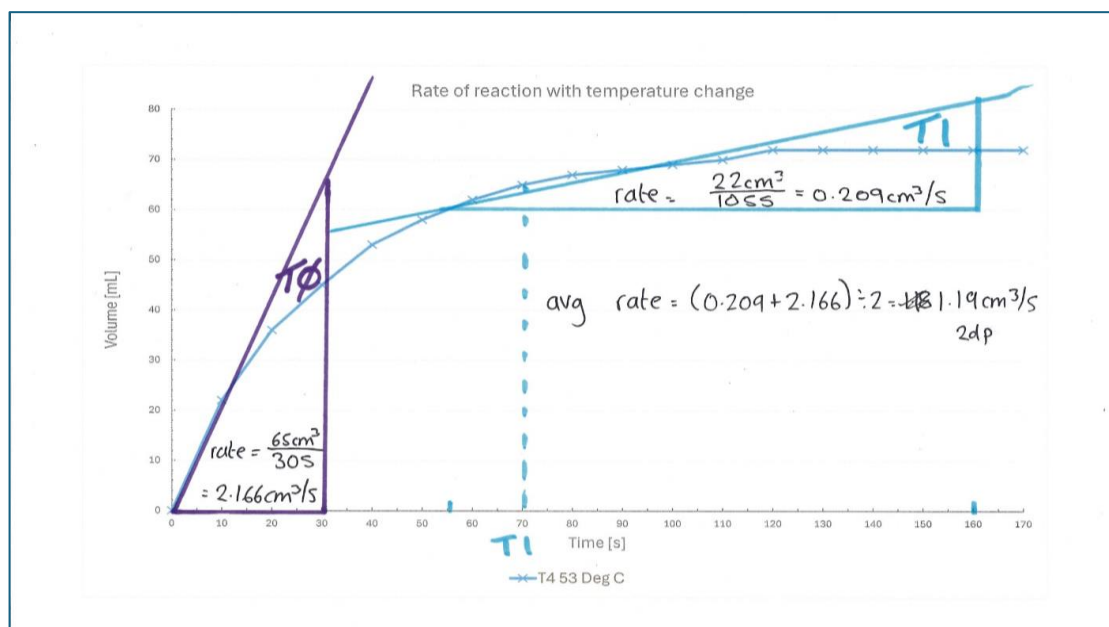


Figure 10 Experiment 2, 1.0 Mol at 53°C

Experiment 2			
Temp [°C]	Time to dissolve Mg [s]	Average Rate [cm^3/s]	Instantaneous Rate @ $t=0$ [cm^3/s]
24.00	170	1.00	1.50
33.00	80	1.16	2.00
43.00	150	1.11	1.87
53.00	120	1.19	2.17

Table 2 Summary of Results for Experiment 2

Experiment 2 Key Findings

The results shown from Experiment 2 support the hypothesis that the rate of reaction increases with temperature. An anomaly however was spotted on T3 which has both a lower average reaction rate and also instantaneous rate which was not expected. This could be potentially caused by error whilst measuring the quantity and / or molarity (concentration) of the HCl solution, or possibly by using an incorrect length of magnesium strip. Although the average rates did not significantly change with temperature, the instantaneous rates $t=0$ did significantly increase with temperatures T1, T2 & T4.

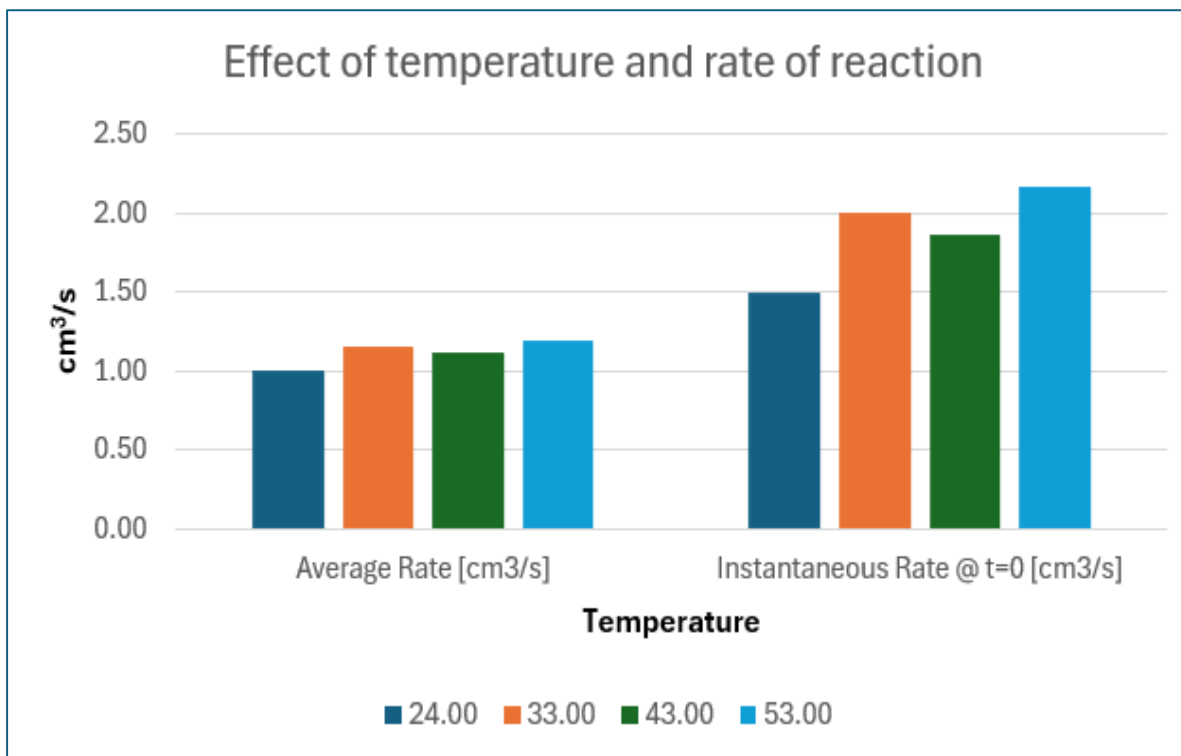


Figure 11 Summary of Results for Experiment 2

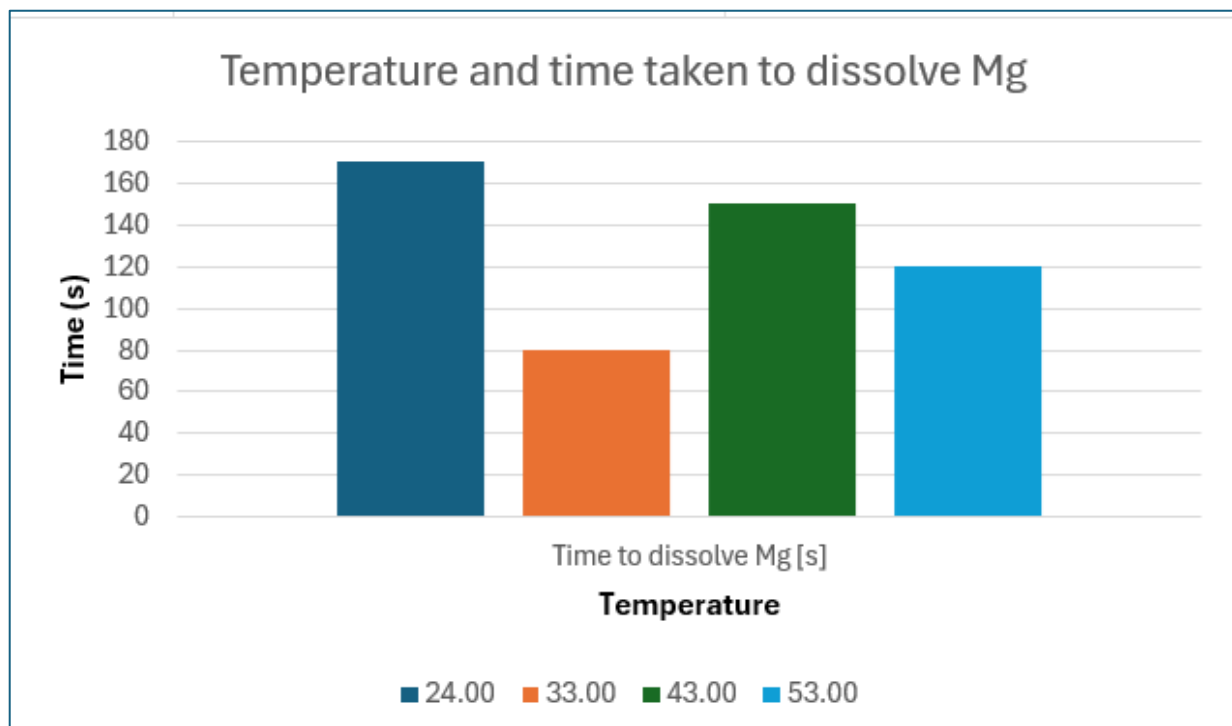


Figure 13 Summary of Results for Experiment 2

Experiment 3 – Effect of Surface Area on Rate of Reaction

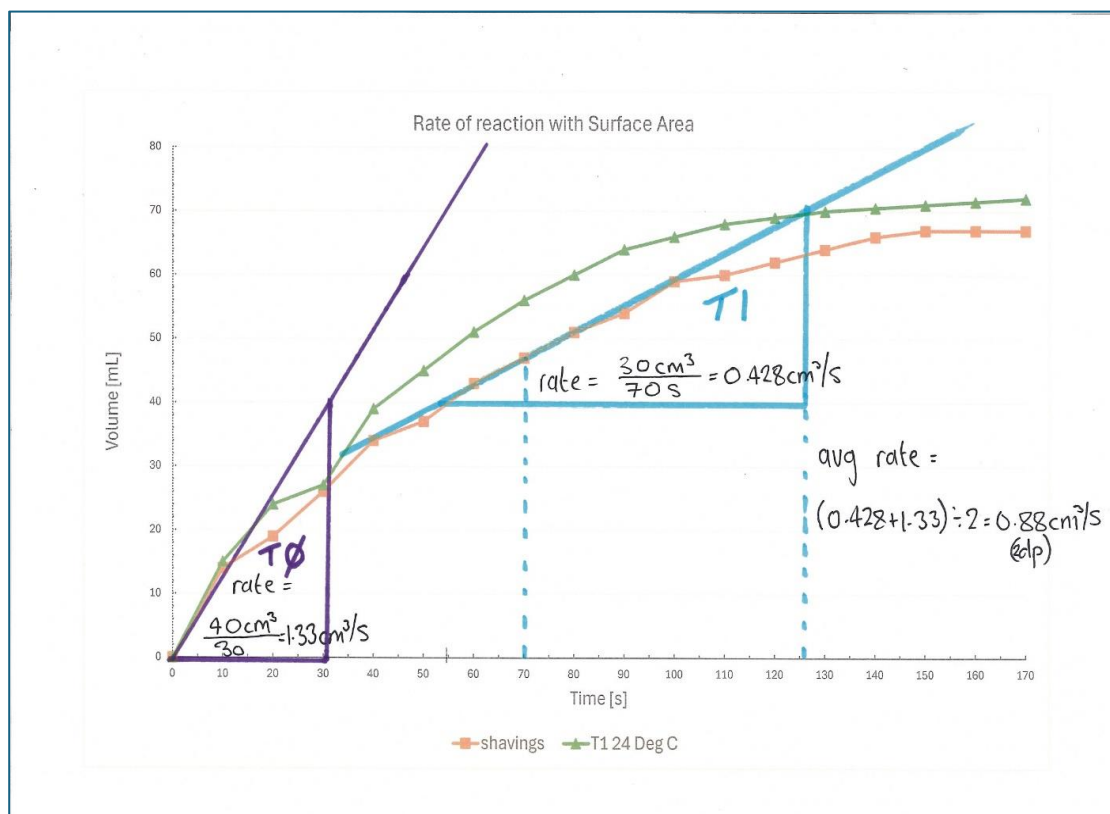


Figure 14 Experiment 3, Turnings/Shavings vs Strip

Experiment 3			
Surface Area [23C 1Mol]	Time to dissolve Mg [s]	Average Rate [cm^3/s]	Instantaneous Rate @ $t=0$ [cm^3/s]
Strip	170	1.00	1.50
Turnings	150	0.88	1.33

Table 3 Summary of Results of Experiment 3

Experiment 3 Key Findings

Table 3 compares data from Experiment 1 dissolving a strip of magnesium in solution (1Mol at 24°C) with the same mass of magnesium but as turnings. The results are shown in Figure 14 and a summary of findings in Table 3. The data shows that the rates of reaction both instantaneous and average are lower with the turnings than the magnesium strip. The graph in Figure 14 also shows that the total volume of gas

4.0 DISCUSSION

Interpretation and Significance of Results & Potential Sources of Experimental Error

The volume of hydrogen measured was dependant on the speed of which the bung could be sealed in the test tube immediately after the magnesium was added once the reaction began. This could cause very small errors in initial gas production readings.

Secondly the molarity (concentration) of the HCl was diluted using a measured quantity of distilled water. Although great care was taken to accurately measure the volumes of HCl and distilled water, no tests were performed to calculate the actual molarity (concentration) of the HCl solution.

The mass of magnesium was to the nearest gram and as the volume of solution was small, small variations in the mass could result in relatively larger or smaller volumes of hydrogen gas which could skew the experimental results.

5.0 CONCLUSIONS

The hypothesis of Experiment 1 was that the rate of reaction increases with increased molarity (concentration) of HCl. The results shown in Figures 1,2,3 and 4 and the summary in Table 1 support the hypothesis that as molarity (concentration) increases, so does the rate of reaction.

Results from Experiment 2 show that the temperature of the HCl solution is proportional to the rate of reaction. Whilst the overall quantity of hydrogen produced did not vary much, the initial and average rate of reaction increased with temperature therefore the hypothesis for Experiment 2 was supported by the experiment.

Experiment 3 did not show expected results. When the reaction using turnings was compared to Experiment 1, using the same molarity (concentration) and temperature of HCl, with a strip of magnesium, the reaction was slower with the turnings. The results do not support the hypothesis that an increase in surface area would cause an increase in the reaction speed [1].

Having researched into the probable causes for this, it is concluded that although the surface area was increased, the surface area of oxidation of the magnesium was also increased [2]. As a result, the increased area of oxidation created a barrier between the HCl and the Magnesium which slowed the initial rate of reaction. As shown in Figure 14, the volume of hydrogen produced using the turnings was also reduced. This was because the increased surface area which was oxidised meaning that there was less pure magnesium to react with the HCl.

Future Research based on Findings

The results show that surface area is not always a good thing when it comes to increasing the rate of reaction, especially if the surface has been exposed to oxygen and become oxidised. Research into alternative ways to carry out this experiment would be beneficial

to improving the accuracy of the results as well as exploring the theory of the role of oxidation in the chemical reaction.

6.0 REFERENCES

- [1] J. Brennan, "How to Explain What Happens When We Burn Magnesium Metal," *Sciencing*, Apr. 25, 2017. <https://sciencing.com/explain-happens-burn-magnesium-metal-8206877.html> (accessed May 21, 2024).
- [2] J. E. Brady and G. E. Humiston, *General Chemistry*. John Wiley & Sons, 1980.
- [3] Khan Academy, "Dalton's Atomic Theory," *Khan Academy*, Jan. 20, 2024. <https://www.khanacademy.org/science/chemistry/electronic-structure-of-atoms/history-of-atomic-structure/a/daltons-atomic-theory-version-2> (accessed May 21, 2024).
- [4] Rocket Chronicles, "High Power Rocket Motor Basics." Accessed: May 21, 2024. [Online]. Available: <https://www.spaceportrocketry.org/pdfs/Rockets%20for%20Schools%20HPR%20Motor%20Basics.pdf#:~:text=The%20propellant%20is%20burned%20to%20provide%20high%20volumes>

7.0 APPENDIX A – Materials

Materials Required for All 3 Experiments

- Long test tube with arm
- Dry gas syringe (do not allow to become wet)
- Rubber bung and delivery tube to fit the test tube
- Two large beakers for the water bath and reaction mixture
- Pipette and pipette filler
- Two clamp stands
- Test tube holder
- Thermometer
- Ring stand, Bunsen burner, clay triangle
- Stopwatch
- Graph paper, pencil and ruler
- Weighing scales

Chemicals

- Magnesium ribbon, cut into 4cm length and weighed
- Magnesium turnings (same mass as ribbon)
- Hydrochloric acid (2.0M and 1.0M)
- DI water

APPENDIX B – Methodology

Methodology Experiment 1

Set Up Using the pipette prepare 4 x 20ml solutions of hydrochloric acid (HCl) with concentrations; 2.0mol, 1.5mol, 1.0mol and 0.5mol.

Attach gas syringe to the first clamp so the syringe is in a horizontal position, and attach the test tube to the second clamp in a vertical position as shown in Figure B1.

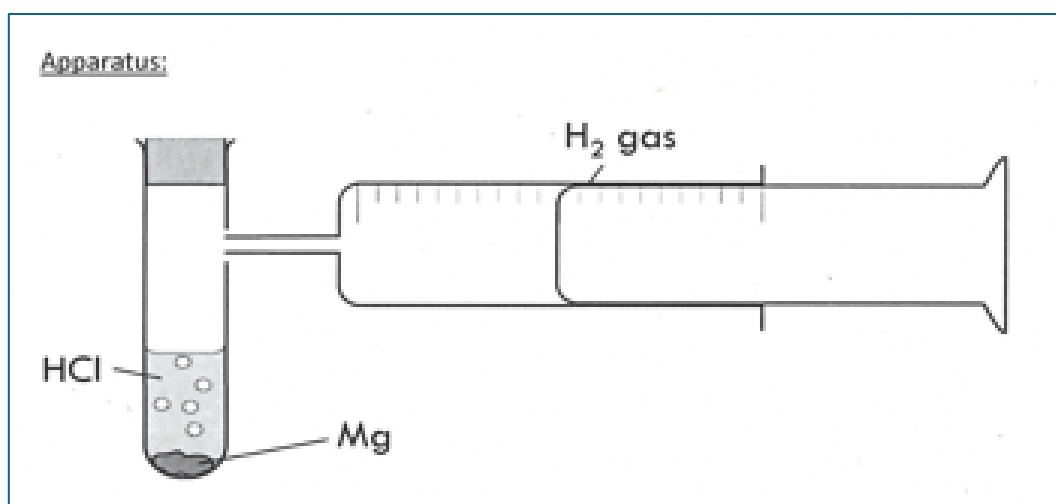


Figure B1 – set up for collection of hydrogen

- Step 1 Pour the 1.0mol concentration of HCl into the test tube and record its temperature.
- Step 2 Ensure the gas syringe is set to zero.
- Step 3 Add a 4cm strip of magnesium to the test tube and replace the bung quickly. Start the stopwatch immediately.
- Step 4 Record the volume of hydrogen gas collected in the syringe at 10 second intervals until no further gas is produced.
- Step 5 Discard the mixture into the waste beaker and reset the syringe to zero.
- Step 6 Repeat the experiments with each of the remaining concentrations, without recording the temperature.
- Step 7 Plot a graph to show the volume of hydrogen produced vs time for each concentration (molarity).

- Step 8 Calculate the average rate of reaction between $T=0s$ and $T=t_1s$ (to be selected and kept consistent for all calculations) as shown in Figure B2.
- Step 9 Calculate the instantaneous rate of reaction $T=0s$ as shown in Figure B2.

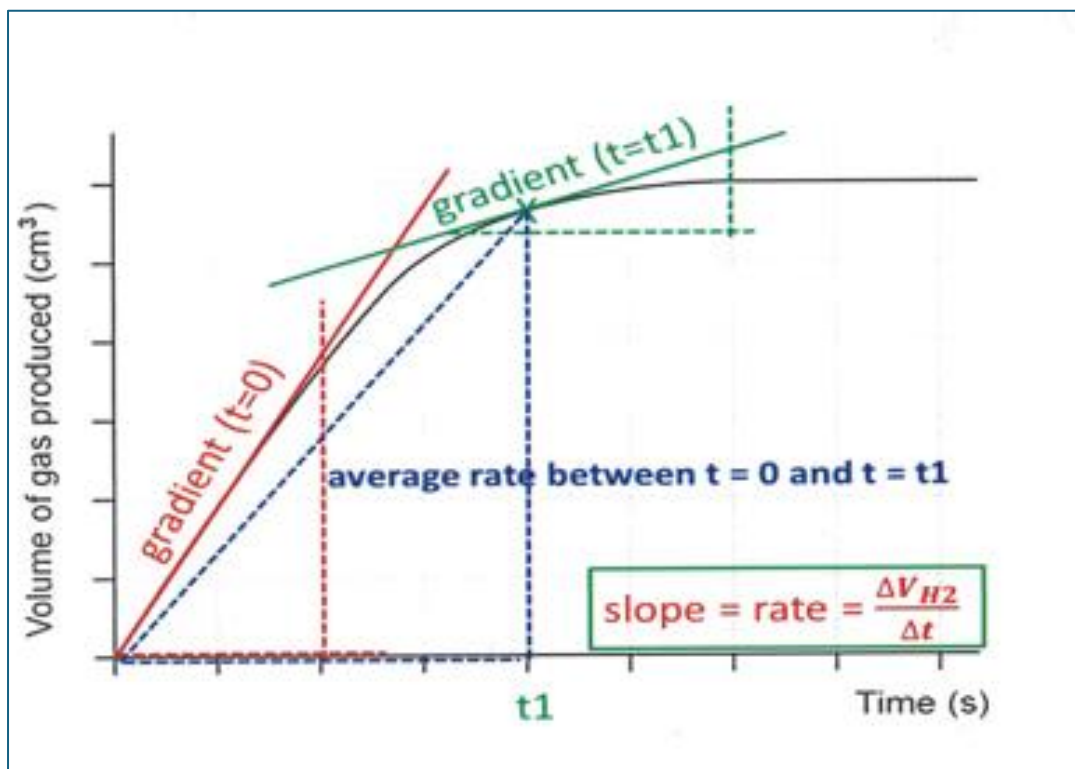


Figure B2 – Calculation of Reaction Rates

Methodology Experiment 2

Set Up Using the pipette prepare 3 x 20ml solutions of hydrochloric acid (HCl) with 1.0mol concentration.

Attach gas syringe to the first clamp so the syringe is in a horizontal position, and attach the test tube to the second clamp in a vertical position as shown in Figure B1 above. Ensure the gas syringe is set to zero.

Set up the water bath on the tripod and clay triangle above a Bunsen burner.

- Step 1 Light the Bunsen burner and heat the water bath until the temperature increases by approximately 10°C above the temperature of the 1.0mol HCl recorded for Experiment 1.
- Step 2 Place a thermometer in the first test tube with 1.0mol HCl in the water bath and wait for the temperature of the HCL to increase to 10°C above the temperature recorded in Experiment 1.
- Step 3 Pour the warm HCl into the vertical test tube and add a 4cm piece of magnesium and replace the bung quickly. Start the stopwatch immediately.
- Step 4 Record the volume of hydrogen gas collected in the syringe at 10 second intervals until no further gas is produced.

- Step 5 Discard the mixture into the waste beaker and reset the syringe to zero.
- Step 6 Repeat the experiments with each of the remaining samples increasing the temperature by approximately 10°C for each experiment.
- Step 7 Plot a graph to show the volume of hydrogen produced vs time for each temperature.
- Step 8 Calculate the average rate of reaction between $T=0s$ and $T=t_1s$ (to be selected and kept consistent for all calculations)
- Step 9 Calculate the instantaneous rate of reaction $T=0s$.

Methodology Experiment 3

Set Up Set Up Prepare 1 x 20ml solution of hydrochloric acid (HCl) with 1.0mol at room temperature.

Attach gas syringe to the first clamp so the syringe is in a horizontal position, and attach the test tube to the second clamp in a vertical position as shown in Figure B1 above.

- Step 1 Pour the 1.0mol concentration of HCl into the test tube.
- Step 2 Ensure the gas syringe is set to zero.
- Step 3 Add the turnings of magnesium to the test tube and replace the bung quickly. Start the stopwatch immediately.
- Step 4 Record the volume of hydrogen gas collected in the syringe at 10 second intervals until no further gas is produced.
- Step 5 Plot a graph to show the volume of hydrogen produced vs time.
- Step 6 Calculate the average rate of reaction between $T=0s$ and $T=t_1s$ (to be selected and kept consistent for all calculations)
- Step 7 Calculate the instantaneous rate of reaction $T=0s$.