

Module - ACDB4011 – L4 Chemistry

Rate of Reaction



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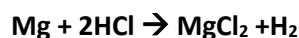
INTRODUCTION

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Purpose: The purpose of this experiment is to investigate the effect that temperature, surface area of the reactant and concentration of an acid has on the rate of reaction, by analysing the speed in which the reactant dissolves into the acid and measuring the volume of gas being released, allowing the application of chemical kinetics within a real life setting rather than purely theoretical. In this case, I will be using Magnesium (Mg) Metal as the reactant, with hydrochloric (HCl) acid as the acid it dissolves in.

The balanced chemical equation for the reaction between Mg and HCl is:



Hypothesis: The Hypothesis of this experiment is to see that when the concentration of the acid, temperature of the acid and surface area of the magnesium increase, the rate of reaction will also increase, showing a directly proportional link between these factors and the rate of reaction. This will be measured via identifying and recording how long the magnesium takes to dissolve, whilst recording how much H₂ gas is released from the reaction over time.

The Variables used can be seen below:

Independent Variables:

- Concentration of the HCl.
 - o Adjusted within the experiment from 0.5M to 2.0M
- Temperature of the HCl.
 - o Adjusted to see impact on the rate of reaction.
- Surface Area of the Magnesium
 - o Changed from an Mg strip, to Mg Fillings.

Dependent Variables:

- The amount of time taken for the magnesium to fully dissolve.
- The volume of H₂ gas produced over the course of the experiment.

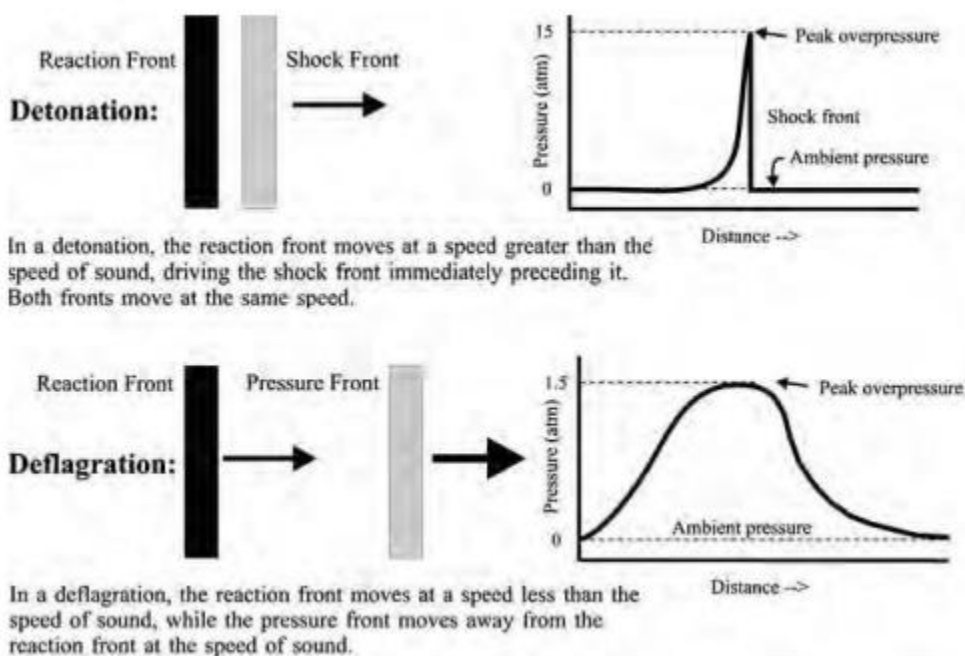
Control Variables:

- The volume of HCl used within each experiment.
- The mass of Magnesium used within each experiment.

THEORY:

- Examples of applications/importance of the kinetics in OME
 - Insensitive Munitions – Using the rate of reaction to categorise how sensitive a munition is.
 - High Explosive fill – Using the rate of reaction to choose a material to be used as a HE.
 - Propellant - Using the rate of reaction to choose a material that will deflagrate and produce a propellant desired effect.

The study of reaction rates within chemistry, known as the study of kinetics, is essential in the field of OME for applications within the field ranging from designing a propellant, to assessing the safety of an explosive. At a very basic level, it is extremely important to understand the kinetics of a material to understand how the OME compound either explodes or deflagrates. These reaction curves can be seen below.



[1] Figure 1: Detonation vs Deflagration

As you can see, these types of reactions are very different. detonations start very calm, before reaching the materials activation energy, and extremely suddenly, increasing in pressure before releasing this energy as an 'explosion' and suddenly dropping down to its baseline. Whereas Deflagration reactions happen by a slow increase and slow decrease, causing a safer reaction due to the energy being released by the exothermic reaction taking place being much calmer and controlled. Understanding this, allows you to advance further into kinetic chemistry, where you can use this information as a starting point to then classify the sensitivity of a munition and how stable it is. I.E, an unstable OME compound, would

have an extremely fast rate of reaction, with a low activation energy, causing a detonation at stable conditions. This would mean that it is extremely dangerous to handle and store, which is essential to understand so that you don't put unsuspecting people in harm's way or have any catastrophic events.

The Arrhenius Equation, $k = Ae^{-E/RT}$, 'can be used to show the effect that a change in temperature has on the rate constant, k , and thus on the overall rate of the reaction' [2]. It allows you to use an equation to predict the stability of a compound and how the variables within the equation, for example the change in temperature, can make the rate constant change, which means that the reaction will happen faster. T

The Equations used within this report can be seen below:

- Moles = Molecular weight / Mass = ($n = MM / m$)
- Rate of Reaction = Change in Volume of H_2 / Change in time = ($\Delta V_{H_2} / \Delta t$)
- Moles = Molarity x Volume (l)

RESULTS

I. EFFECT OF HCL CONCENTRATION ON THE SPEED OF H₂ GAS EVOLUTION

Moles of reactants:

Moles = Mass (g) / Molar Mass (g/mol)

- Mg Moles = $0.082/24.305 = 0.003374$

Moles = Molarity x Volume (l)

- Convert 4 volumes into (l)
 - o 20ml = 0.02l
 - o 15ml = 0.015l
 - o 10ml = 0.01l
 - o 5ml = 0.005l
- Put into Moles Equation.
 - o $2.0M \times 0.020l = 0.04 \text{ Mol/HL}$
 - o $2.0M \times 0.015l = 0.03 \text{ Mol/HL}$
 - o $2.0M \times 0.01l = 0.02 \text{ Mol/HL}$
 - o $2.0M \times 0.005l = 0.01 \text{ Mol/HL}$

Mg		2 M HCL	
Mass	moles	volume	moles
0.082g	0.003374 Mol	20 ml	0.04 Mol/HL
		15 ml	0.03 Mol/HL
		10 ml	0.02 Mol/HL
		5 ml	0.01 Mol/HL

Results Overview

0.5 M HCL	
time, s	V H ₂ , mL
10	2
20	4
30	5
40	7
50	9.5
60	11
70	12.5
80	14
90	15.5
100	17
110	19
120	19.5
130	21
140	22.5
150	24
160	25
170	26.5

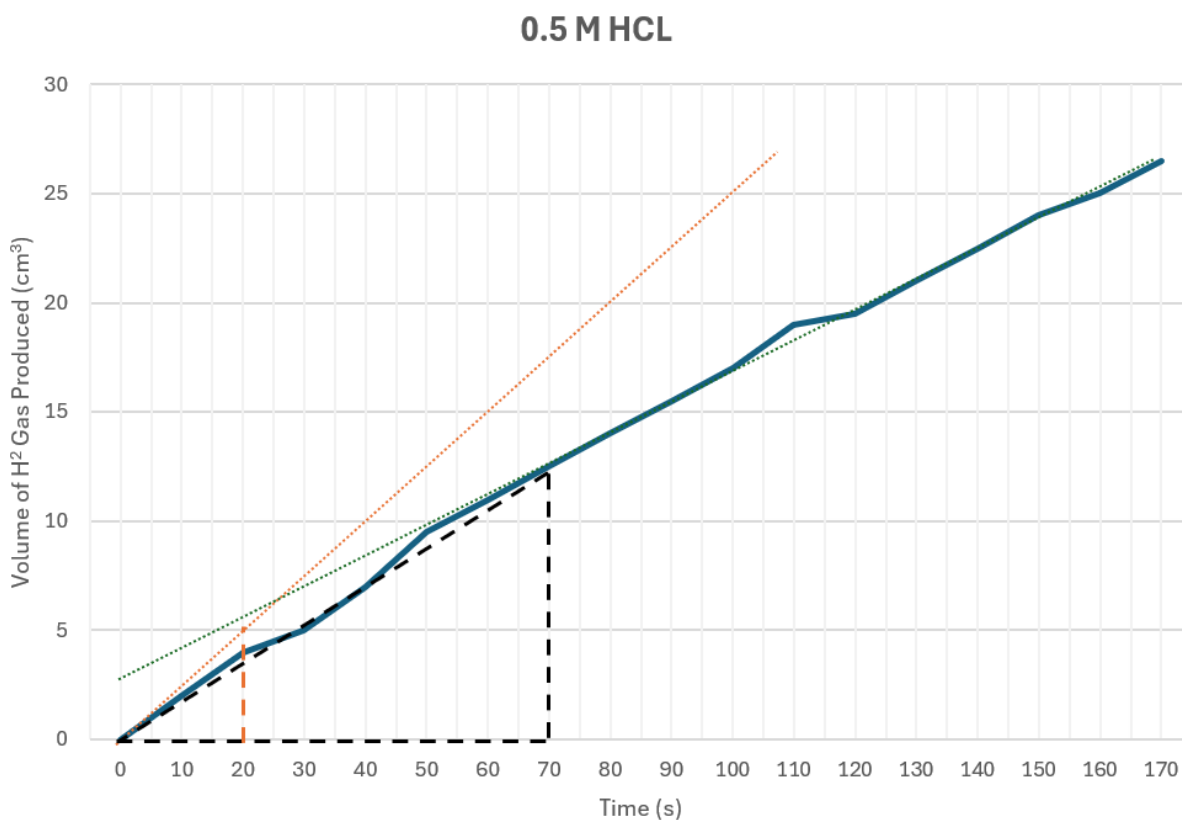
1.0 M HCl	
T 1 = 27°C	
time, s	V H ₂ , mL
10	18
20	30
30	36
40	45
50	52
60	56
70	66
80	68
90	72
100	73
110	75
120	76
130	77
140	77.5
150	78
160	79
170	80

1.5 M HCl	
time, s	V H ₂ , mL
10	20
20	38
30	46
40	56
50	62
60	65
70	69
80	74
90	75
100	78
110	82
120	86
130	86
140	86
150	86
160	86
170	86

2.0 M HCl	
time, s	V H ₂ , mL
10	36
20	38
30	66
40	70
50	75
60	78
70	84
80	84
90	84.5
100	85
110	85
120	85
130	85
140	85
150	85
160	85
170	85

Result 1 – 0.5 M HCL

Graphical Plot

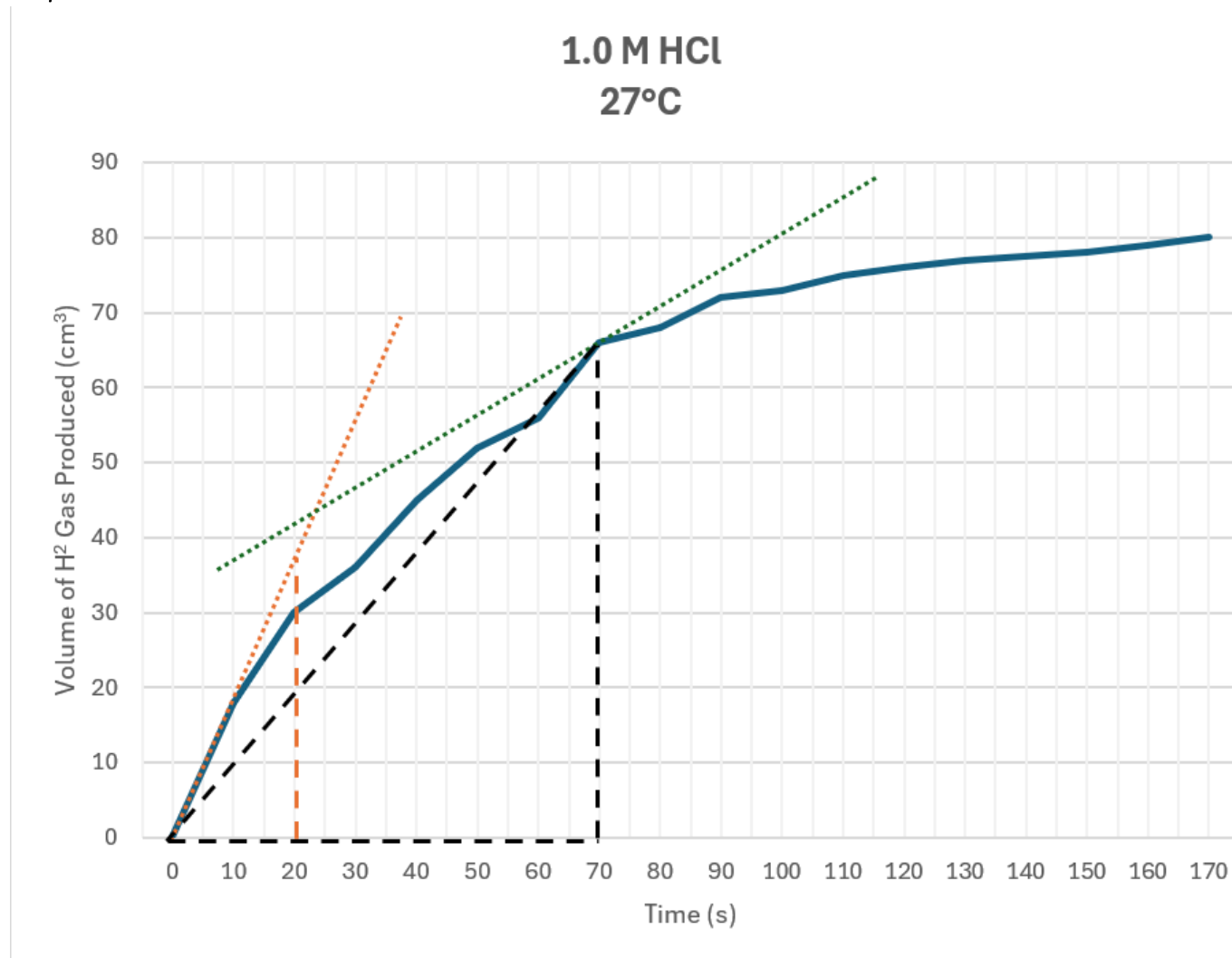


A) Average Rate of Reaction = $12.5/70 = 0.17857 \text{ cm}^3/\text{s}$ – Values taken from the black Triangle of the graph.

B) Instantaneous Rate of Reaction = $5/20 = 0.25 \text{ cm}^3/\text{s}$ – Values taken from the orange Triangle of the graph.

Result 2 – 1.0 M HCL

Graphical Plot

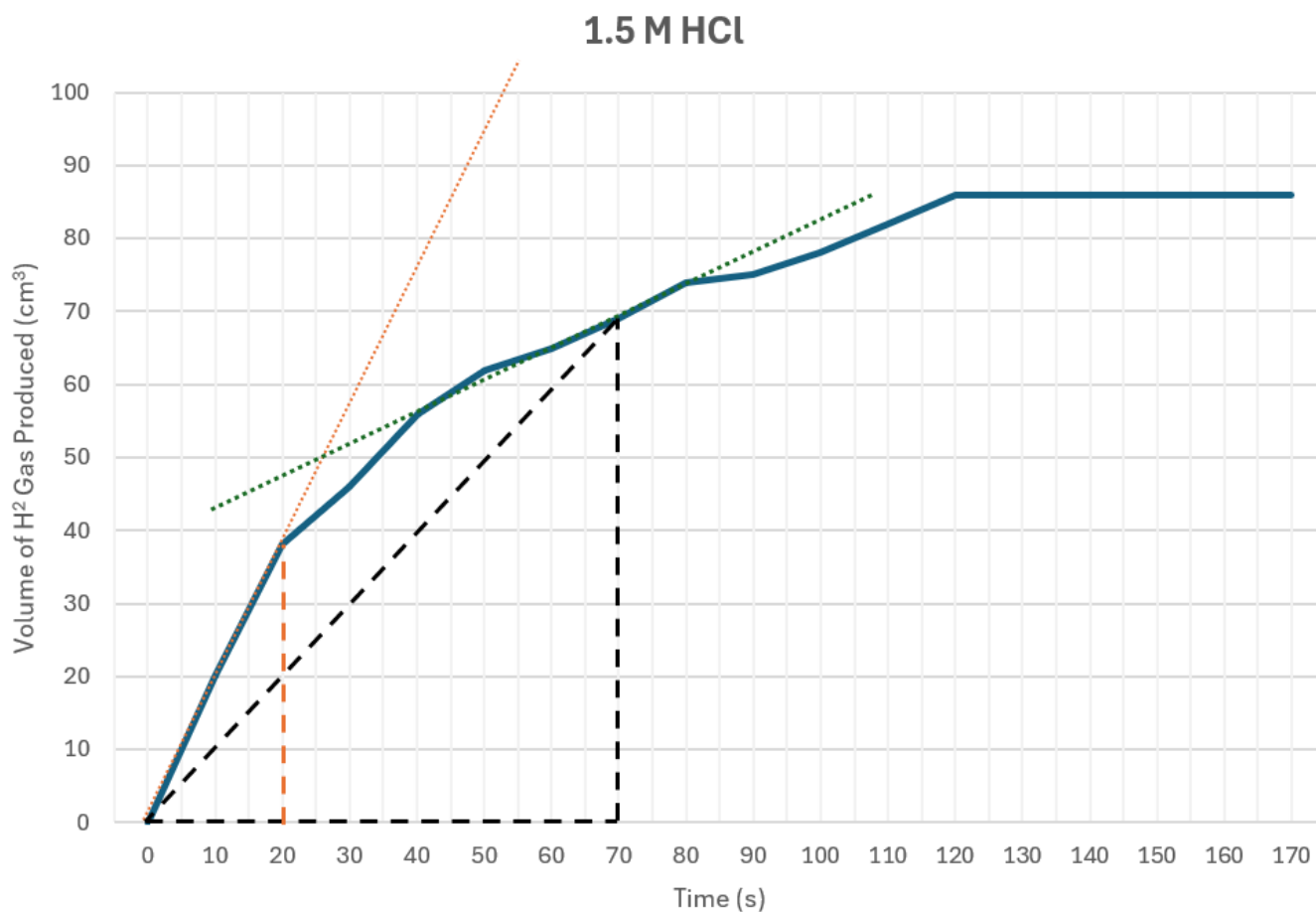


A) Average Rate of Reaction = $66/70 = 0.95 \text{ cm}^3/\text{s}$ – Values taken from the black Triangle of the graph.

B) Instantaneous Rate of Reaction = $37/20 = 1.85 \text{ cm}^3/\text{s}$ - Values taken from the orange Triangle of the graph.

Result 3 – 1.5 M HCL

Graphical Plot

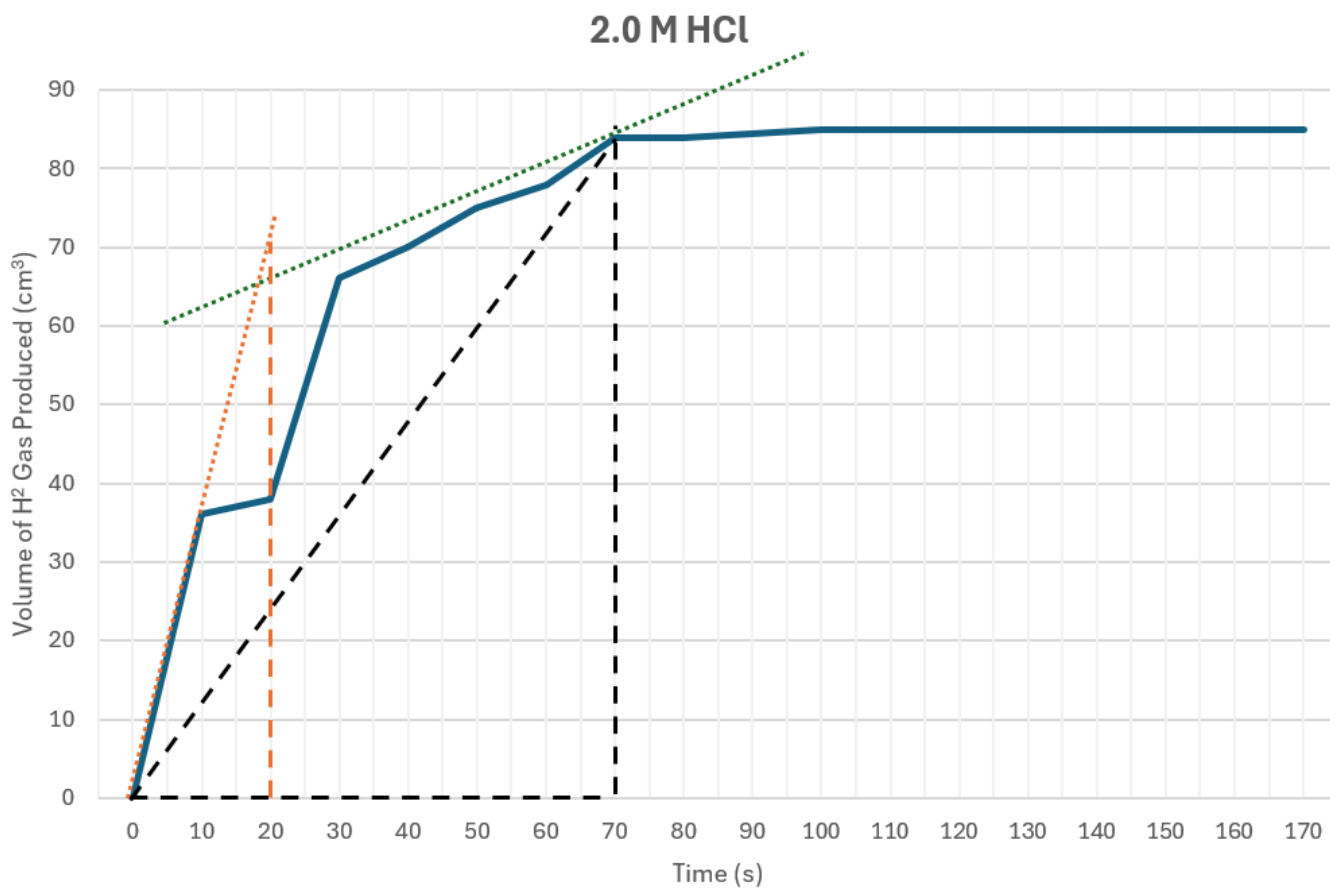


A) Average Rate of Reaction = $69/70 = 0.99 \text{ cm}^3/\text{s}$ – Values taken from the black Triangle of the graph.

B) Instantaneous Rate of Reaction = $40/20 = 2 \text{ cm}^3/\text{s}$ - Values taken from the orange Triangle of the graph.

Result 4 – 2.0 M HCL

Graphical Plot



A) Average Rate of Reaction = $84/70 = 1.2 \text{ cm}^3/\text{s}$ – Values taken from the black Triangle of the graph.

B) Instantaneous Rate of Reaction = $70/20 = 3.5 \text{ cm}^3/\text{s}$ - Values taken from the orange Triangle of the graph.

Results Comparison

[HCl]	time to dissolve Mg	average rate	instantaneous rate at
			t = 0
0.5 M	210	0.17857 cm ³ /s	0.25 cm ³ /s
1.0 M	180	0.95 cm ³ /s	1.85 cm ³ /s
1.5 M	140	0.99 cm ³ /s	2 cm ³ /s
2.0 M	120	1.2 cm ³ /s	3.5 cm ³ /s

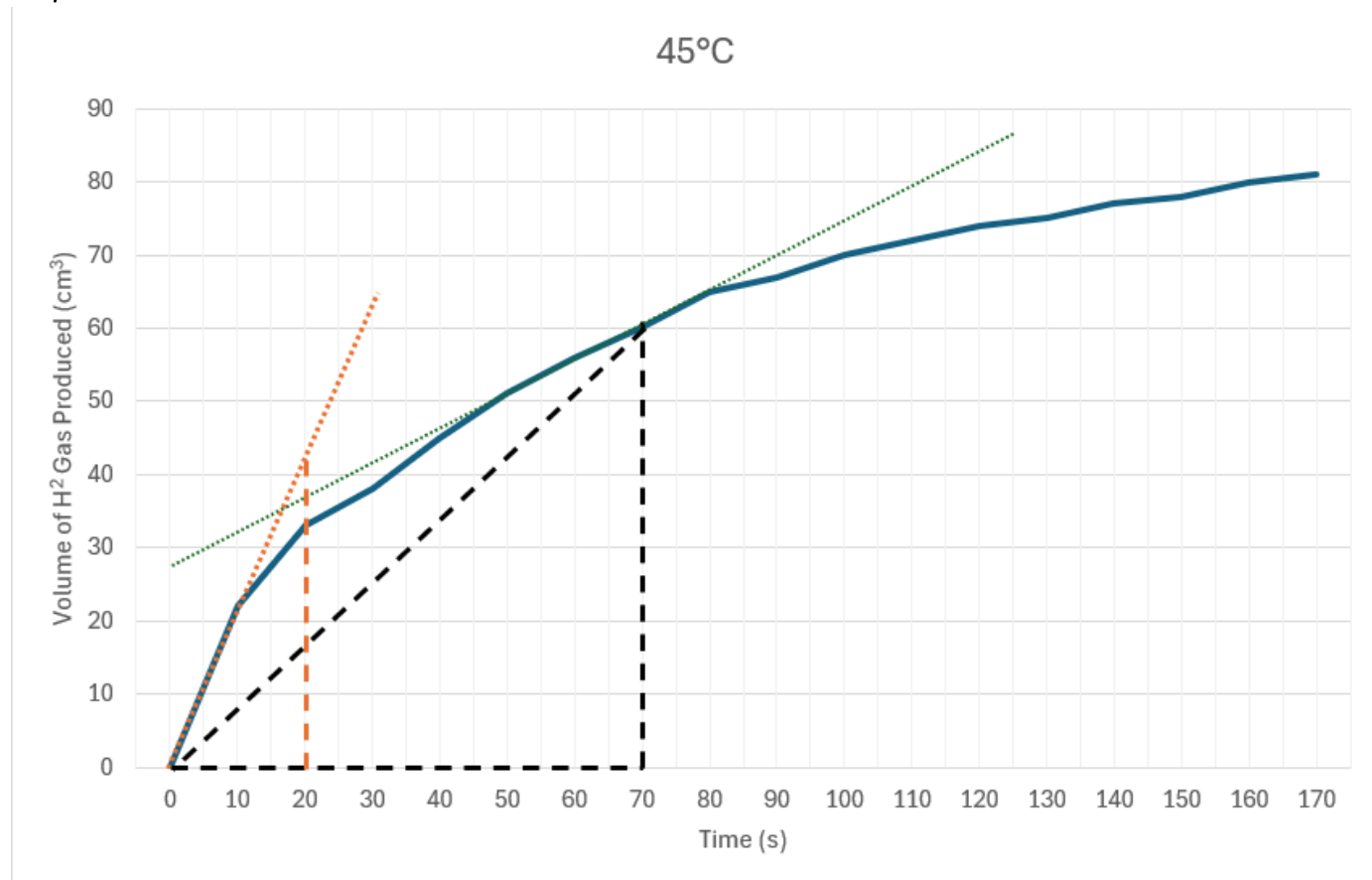
II. EFFECT OF TEMPERATURE ON THE SPEED OF H₂ GAS EVOLUTION

Results Overview

T 2 = 45°C		T 3 = 55°C	
time, s	V H ₂ , mL	time, s	V H ₂ , mL
10	22	10	22
20	33	20	30
30	38	30	37
40	45	40	43
50	51	50	47
60	56	60	53
70	60	70	58
80	65	80	60
90	67	90	63
100	70	100	65
110	72	110	66
120	74	120	67
130	75	130	68
140	77	140	70
150	78	150	71
160	80	160	72
170	81	170	75

Result 1 – 45°C Test

Graphical Plot

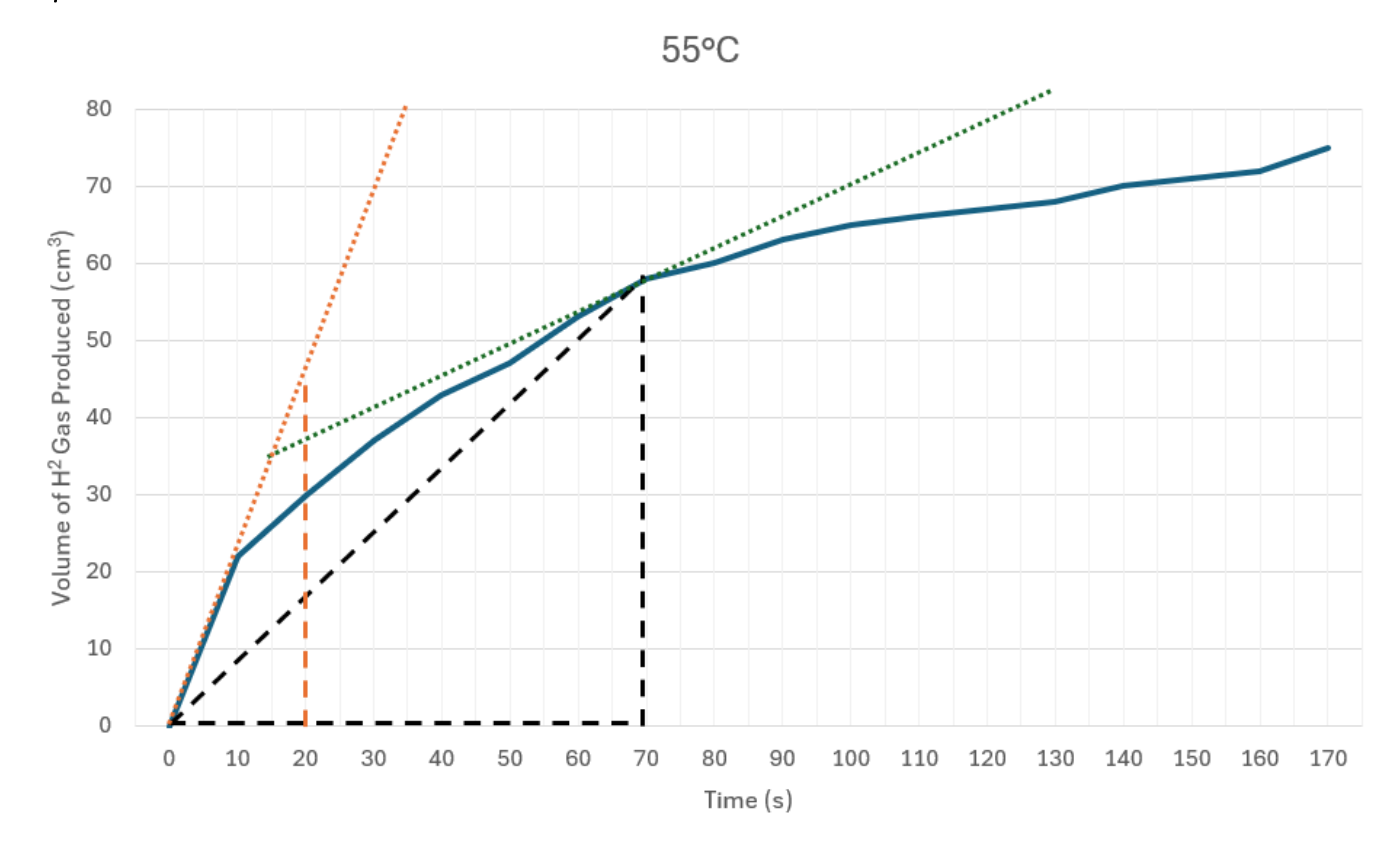


A) Average Rate of Reaction = $60/70 = 0.86 \text{ cm}^3/\text{s}$ – Values taken from the black Triangle of the graph.

B) Instantaneous Rate of Reaction = $42/20 = 2.1 \text{ cm}^3/\text{s}$ - Values taken from the orange Triangle of the graph.

Result 2 – 55°C Test

Graphical Plot



A) Average Rate of Reaction = $58/70 = 0.83 \text{ cm}^3/\text{s}$ – Values taken from the black Triangle of the graph.

B) Instantaneous Rate of Reaction = $45/20 = 2.25 \text{ cm}^3/\text{s}$ - Values taken from the orange Triangle of the graph.

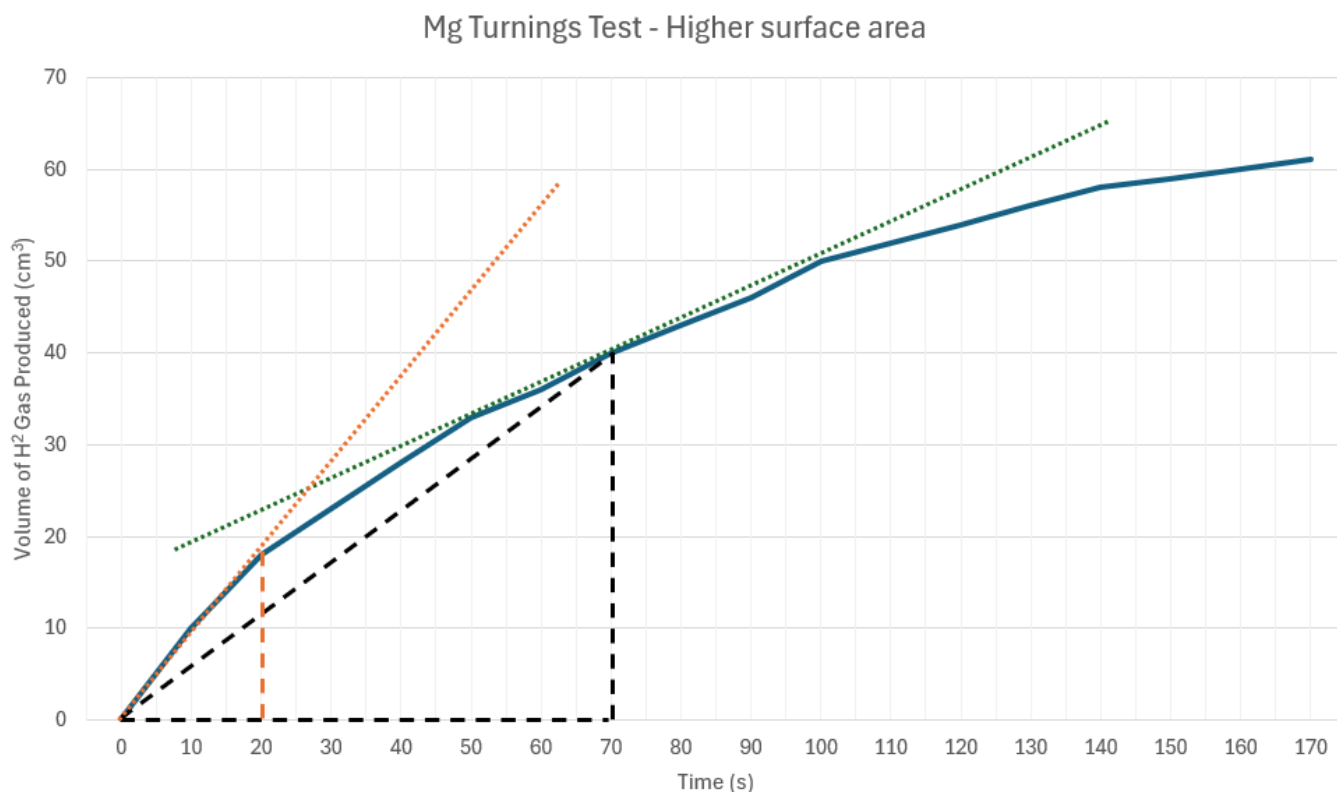
Results Comparison

temperature	time to dissolve Mg	average rate	instantaneous rate at
			t = 0
T 1 - 27°C	180s	1.5 cm ³ /s	1.85 cm ³ /s
T 2 - 45°C	170s	0.86 cm ³ /s	2.1 cm ³ /s
T 3 - 55°C	190s	0.83 cm ³ /s	2.25 cm ³ /s

III. EFFECT OF THE SURFACE AREA OF Mg ON SPEED OF H₂ GAS EVOLUTION

Result 1 – Mg Turnings Test

Graphical Plot



A) Average Rate of Reaction = $40/70 = 0.57 \text{ cm}^3/\text{s}$ – Values taken from the black Triangle of the graph.

B) Instantaneous Rate of Reaction = $18/20 = 0.9 \text{ cm}^3/\text{s}$ - Values taken from the orange Triangle of the graph.

Mg turnings		rate	
mass	moles	average	instantaneous at t = 0
0.082g	0.003374 Mol	$40/70$ $=0.57143$ cm^3/s	$18/20 = 0.9 \text{ cm}^3/\text{s}$

DISCUSSION

EFFECT OF HCl CONCENTRATION ON THE SPEED OF H₂ GAS EVOLUTION

The outcome of this experiment was that as the strength of the HCl concentration was increased, the rate of reaction and instantaneous rate both increased, whilst the time taken for the Mg to dissolve decreased. This shows that when you increase the concentration of the acid, the reaction happens at a faster speed. This result has backed the original hypothesis, which is also agreed online where 'The higher the concentration of reactants, the faster the rate of a reaction will be.' [3] This can be seen on the graphs provided.

EFFECT OF TEMPERATURE ON THE SPEED OF H₂ GAS EVOLUTION

This test was meant to prove the hypothesis stating that an increase in temperature would increase the rate of reaction. This was the case for the 45°C result, which took less time for the Mg to dissolve than that at 27°C. However, the results gathered shown the experiment ran at 55°C, to take the longest to dissolve and have the lowest rate of reaction. However, the instantaneous rates show the expected outcome. Because of this, it appears that after the first 20 seconds or so, there has been a failure within the test, which may have come from the User, unsatisfactory monitoring of the results, which is the most likely scenario since this experiment was rushed to fit inside the stretched time limit, causing one of the tests to not occur. 'An increase in temperature typically increases the rate of reaction. An increase in temperature will raise the average kinetic energy of the reactant molecules.' [4]

EFFECT OF THE SURFACE AREA OF Mg ON SPEED OF H₂ GAS EVOLUTION

Within this, the results are much lower than expected. The increase in surface area would be expected to cause a dramatic increase in the rate of reaction and cause full dissolution of the Mg within the HCl much faster than the Mg Strips would have. This is because 'The more surface area that is available for particles to collide, the faster the reaction will occur.' [5] The results not being what was expected could be down to several experimental results, listed below. However, it could show that temperature potentially plays a more important role than surface area. There could be a potential future experiment to investigate this hypothesis.

EXPERIMENTAL ERRORS

User Error – Within this experiment, full attention to both the rapidly rising gas syringe as well as to the stopwatch and magnesium was required. Because of this, it would have been easy to neglect one or more of these parameters by accident or misread a result. To prevent this from happening in the future, automating the thermometer to record a result every 10 seconds via electronic means could be implemented, albeit an expensive measure. Or, you could have one person monitoring every parameter and implement a timer with sounds for the stopwatch, so rather than watching the clock, you could hear a buzz and then know when to take a measurement.

Oxidation Layer – An oxidation layer would have formed on the magnesium metal, which reduces the rate of reaction by implementing a barrier between the metal and HCl, this barrier would vary from piece to piece adjusting by size of the material or weight. This could be resolved by ensuring all the Magnesium pieces were the same size and exact shape.

Different size materials – Due to being hand cut, the size of the magnesium could have been off slightly, with different sources of error creeping in from all angles such as human error, issues with tooling, the cutting environment and lack of precision being used in a low budget test. This could be rectified by getting the materials brought to size, or by automating the cutting process using technology.

CONCLUSIONS

This paper shows the use of rate of reaction and kinetic chemistry within the world of OME engineering, whilst providing an in-depth experience with detailed results, including signs of error and future ideas to conduct an improved experiment. It has, backed with online data, dived into the three main parameters of rate of reaction, concentration, surface area and temperature and explained all three and how they are modified to obtain the desired results.

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APPENDIX A: CHEMICALS & EQUIPMENT USED:

Chemicals:

- Magnesium 4cm strips, all weighing 0.082g
- Magnesium turnings weighing 0.082g
- Hydrochloric Acid, both 1.0M and 2.0M
- DI water

Equipment:

- Gas syringe
- Integrated Test Tube with Arm
- Rubber Bung with delivery tube that fits the test tube attached to the arm.
- Waste Beaker
- Water Bath Beaker
- Pipette
- Pipette Filler
- 2 Clamp stands
- Test Tube Holder
- Thermometer
- Phone (stopwatch)
- Bunsen Burner with a clay triangle and ring stand

APPENDIX B: PROCEDURE

CONCENTRATION TEST

1. Place 4 beakers side by side
2. Use pipette and HCL to fill each beaker to 20m of Hydrochloric acid, varying the concentrations for each beaker using DI water to make one 0.5 M, 1.0 M, 1.5 M and 2.0 M
3. Pour the lowest concentration into the test tube attached to the arm.
4. Step forward, place a Mg strip into the test tube, press start on your stopwatch and then place the bung back on.
5. Record the amount of Hydrogen gas given off every 10 seconds.
6. Continue this until no more gas is released.
7. Now discard the waste mixture into the waste beaker.
8. Repeat with the other concentrations. However, measure the temperature of the 1.0 M concentration.
9. Plot a graph of the Volume of the gas produced vs the duration of time.
10. Use this to calculate the Average rate of reaction and the instantaneous rate of reaction at $t = 0$

TEMPERATURE TEST

1. Fill 2 test tubes with 20mL of 1.0 M HCL using a pipette.
 - a. Note 3 test tubes is recommended. Within this experiment, time ran out so only 2 was completed.
2. Place your water bath beaker in front of you and fill with water.
3. Heat the water bath to 10 degrees Celsius higher than the recorded temperature of 1 M HCL concentration.
4. Place one of the test tubes into the water bath and wait until the acid reaches the same temperature of the water.
5. Now transfer the heated HCL into the test tube attached to the arm, before putting in a magnesium ribbon, starting the stopwatch and placing the bung back onto the test tube.
6. Record data every 10 seconds and stop once no more gas is released.
7. Now empty waste mixture into the waste beaker.
8. Repeat this for the next test tube of HCL.

9. Plot a graph of the Volume of the gas produced vs the duration of time.
10. Use this to calculate the Average rate of reaction and the instantaneous rate of reaction at $t = 0$

SURFACE AREA TEST

1. Put 20mL of 1.0M HCL into the test tube attached to the arm.
2. Add the Magnesium turnings into the test tube, start the stopwatch and put the bung back on. Do this simultaneously.
3. Record data every 10 seconds until no more gas is released.
4. Plot a graph of the Volume of the gas produced vs the duration of time.
5. Use this to calculate the Average rate of reaction and the instantaneous rate of reaction at $t = 0$