

Kinetics of the reactions between magnesium and acids



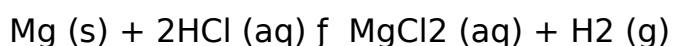
This is due to the dissociation of H ions once the acid is hydrated with water solution to become aqueous. There is only a maximum number of hydrogen atoms in water, hydrochloric acid only consists of one hydrogen atom hence it can only dissociate once producing only 1 H⁺ ion (proton). This proton then bonds with the water molecules to produce a hydronium ion H₃O⁺. This leaves the Cl⁻ in solution. HCl is a strong acid due to its full dissociation in water. This is very unstable (very reactive) as the negative charge is only present on the chloride ion. Moving on to sulphuric acid, it consists of 2 hydrogen atoms resulting in two H⁺ ions (protons) forming, the reaction is highly exothermic. However, the full dissociation of the acid occurs in 2 reactions due to the 2 hydrogen atoms, the first step is $\text{H}_2\text{SO}_4 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{HSO}_4^-$. The hydrogen attached to the sulphate group in the product reacts with water again forming: $\text{HSO}_4^- + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{SO}_4^{2-}$ leaving the sulphate on its own. It is for this reason why sulphuric acid is weaker than hydrochloric acid. The negative charge is initially spread into a hydrogen ion thus making it a weak acid than HCl. And finally ethanoic acid contains four hydrogen atoms, dissociation is very limited and very few hydrogen atoms are dissociated this is all due to the acid itself containing 4 hydrogen; again it forms a hydronium ion. In this case the negative charge is spread across the -COO⁻ group, which makes this acid a weak acid. Carboxylic acids themselves are known as weak acids due to this dissociation.

From what we know Acid + Metal \rightarrow Salt + hydrogen

In all three reactions the magnesium ribbon will dissolve in the acid producing hydrogen gas. Magnesium will react with these acids because it is

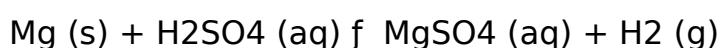
higher in the reactivity series than hydrogen, thus the magnesium will replace the hydrogen in the acids forming the magnesium salts.

Reaction between magnesium and hydrochloric acid:



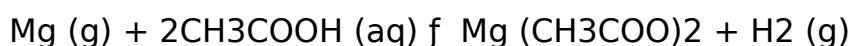
Magnesium + hydrochloric acid \rightarrow magnesium chloride + hydrogen

Reaction between magnesium and sulphuric acid:



Magnesium + sulphuric acid \rightarrow magnesium sulphate + hydrogen

Reaction between magnesium and ethanoic acid



Magnesium + ethanoic acid \rightarrow magnesium ethanoate + hydrogen

In order for magnesium to react with an acid it, both substances must collide. The collision theory simply states that at higher concentrations there is more chance of the two species to collide where as if concentration is low, collision is less frequent. This is because at higher concentration there are more particles in a limited space, therefore under greater pressure. For a collision, hence a reaction, to occur the species must collide with the precise minimum energy called the activation energy of the reaction. If, for example, the species collide with less energy than the activation energy there would be no reaction, they would simply bounce off each other.

The changing of concentrations usually changes the rate of the reaction; from this I was able to identify the order of reactions for each experiment. A rate equation shows the effect of changing concentrations mathematically, there cannot be known unless a reaction is carried out and data is taken.

The rate of reaction is defined as the rate at which reactants are converted into products. If we use the example of the reaction between magnesium and hydrochloric acid, the rate reaction is the rate at which magnesium chloride and hydrogen is made and at the same time the rate at which magnesium and hydrochloric acid is used up.

In all my reactions I measured the rate of reactions by measuring the volumes of gases (hydrogen) released. The hydrogen gas was collected in a gas syringe; the volume produced can be used to determine the rate of reaction. (see picture below)

After obtaining a sufficient amount of data I was able to find the initial rate of the reactions by plotting graphs and calculating the ignition gradient (time= 0) simply by this mathematical equation:

Change in y-axis \div change in x-axis

This is because the initial rate = the initial gradient

After carrying this out I was able to plot a rate-concentration graph. The shape of this graph concluded the order of the reaction.

A rate equation cannot be predicted from a reaction from its balanced equation. The only way to find a rate equation for a reaction is by doing experiments to find the effect of varying the concentrations of reactions.

The general reaction



Therefore the rate equation would be:

$$\text{Rate} = k[A]^m[B]^n$$

Rate is measured in $\text{mol dm}^{-3}\text{s}^{-1}$ and the concentrations [A] and [B] are measured in mol dm^{-3} .

Key:

[A] and [B] are concentrations of reactants A and B

k is the rate constant for the reaction at a particular temperature

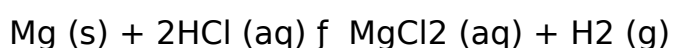
m and n are the orders of reactions with respect to A and B respectively

(n+m) is the overall order of the reaction

In my investigation I explored the change in rate with varying temperature.

Predicting the order of reactions (k= constant e. g. temperature):

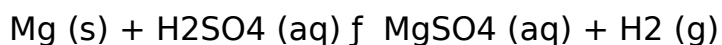
Using hydrochloric acid



$$\text{Rate} = k[\text{Mg}][\text{HCl}]^2$$

The reaction is:

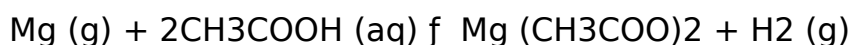
Using sulphuric acid



$$\text{Rate} = k [\text{Mg}][\text{H}_2\text{SO}_4]$$

The reaction is:

Using ethanoic acid



$$\text{Rate} = k [\text{Mg}][\text{CH}_3\text{COOH}]^2$$

The reaction is:

Acid/metal background information:

Acid/metal

Information

Magnesium (Mg)

Atomic number is 12

Alkaline earth metal

Group 2

Period 3

Block s

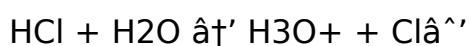
Solid phase

Hydrochloric acid (HCl)

Hydrogen chloride solution

Highly corrosive

In aqueous solution it contains H_3O^+ and Cl^- ions. A H^+ ion from the acid combines with H_2O and forms the H_3O^+ hydronium ion.



Molar mass is 36.46

Phase: colourless liquid at room temperature

Melting point is 247K

Boiling point is 383K

Sulphuric acid (H_2SO_4)

Strong mineral acid

Soluble in water at all concentrations

Molar mass is 98.08

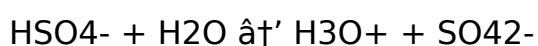
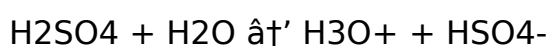
Phase: colourless and odorless liquid

Melting point is 283K

Boiling point is 337K

Created in a manufacture process known as the " contract process (DCDA)"

Creating aqueous sulphuric acid is highly exothermic therefore water is always added to the acid rather the vice versa. Reactions are:



Ethanoic acid (CH_3COOH)

Organic, carboxylic acid

Weak acid (only partially dissociated in aqueous solution)

Known as acetic acid

Involved in vinegar

Colourless liquid

Molar mass is 60. 05

Melting point 290K

Boiling point 391K

The hydrogen atom on the carboxyl group on the acid (-COOH) can be given off as an H⁺ ion (proton) - this give them their acidic character. Again a hydronium ion is formed.

My hypothesis: the rate of reaction for the reaction between hydrochloric acid and magnesium will be greater than that of sulphuric acid and the reaction involving ethanoic acid will be less than both. Increasing the temperature increases the constant k, therefore the rate of reaction will increase, and decreasing the temperature will decrease the rate of reaction as the constant k decreases. The rate of reaction is proportional to the concentration

(Test investigation)

(What else can be measured)

Risk assessment:

(Corrosive) (Irritant) (Toxic) (Flammable) (Explosive)

Name

Hazard

Precautions of handling

Emergency action

Hydrochloric acid

Corrosive - causes burn

Irritant - can irritate the respiratory system

Toxic - inhalation of HCl vapor can cause severe injuries.

Solutions greater of equal to 6.5M are corrosive.

Solutions between 2M and 6.5M are irritant

Avoid heat and flames.

Wear safety glasses or face mask

Wear gloves

Wear lab coat

Eye contact - rinse eyes out immediately and seek medical advice

Hazard area must be isolated

If spilt, ventilate the area. The acid must be neutralized with an alkaline, and then absorbed with a chemical resistant kit and placed in a disposal bag.

Ingestion - wash mouth out, drink plenty of water, and seek medical attention.

Remove contaminated clothing

Sulphuric acid

Corrosive - causes serious burns

Toxic and irritant- harmful by inhalation, ingestion and by skin contact.

Can cause severe burns on skin, chronic exposure can lead to lung damage and lung cancer.

Flammable

Explosive

Solutions stronger or equal to than 1.5M are corrosive.

Solutions between 0.5M and 1.5M are irritant.

Dangerous with water

Wear eye protection (safety goggles).

Take extra care with higher molarities.

Wear gloves to not allow the acid to come into contact with skin.

Whilst diluting, ALWAYS add cold water to the acid, NEVER vice versa. Stir frequently to ensure a mix.

If gets into eyes use an emergency eye wash, flood the eye with the water for about ten minutes.

Wash hands if gets on skin (wear nitrile gloves if supplied).

Clean up area if split, small amounts may be flushed down the sink with a large volume of water following it. Large amounts of spillage must be neutralized with either lime or soda ash before discarding.

Remove contaminated clothing.

<https://assignbuster.com/kinetics-of-the-reactions-between-magnesium-and-acids/>

Ethanoic acid

Flammable

Corrosive

Toxic

Solutions stronger or equal to 4M are corrosive

Solutions between 1.5M and 4M are irritant

Keep away from flames

Wear goggles

Spills can be cleaned up using paper towels, leaving paper towels to dry in fume cupboard before disposal. Proceed by cleaning site with plenty amount of water.

Contact with eyes or skin - wash with large amount of water.

Remove contaminated clothing

Magnesium and its salts

Highly flammable, however difficult to ignite and difficult to extinguish.

Powdered magnesium must not be blown, can ignite in presence of a Bunsen flame.

Contact with mouth - wash mouth out immediately, drink water. Medicine attention must be obtained as side effects such as diarrhoea can show.

Contact with eye - flood eye with water, medical attention must be obtained

Spillage - removed from area, area must be washed.

Equipment used:

Several beakers with the different concentrations of the different acids

Measuring cylinders

Pipettes

Ruler

Pencil

Scissors

100ml conical flasks

Magnesium ribbon

3M of Hydrochloric acid

3M of Sulphuric acid

3M of Ethanoic acid

Water tank

Stopwatch

100cm³ gas syringe, attached to a rubber hose and a hung.

I was provided with only 3M acids, I had to therefore dilute the acids to get the concentrations I needed. I did this by the following steps:

To obtain 2.5M from the 3M acid I:

Disinfected and cleaned several beakers and measuring cylinders

Poured a reasonable amount of the acid into a beaker

Then I poured 90cm³ of the acid into a measuring cylinder and put it into another fresh beaker

Using another cylinder I obtained 75cm³ of cold distilled water

Carefully I poured the cold distilled water to the acid stirring gently allowing a good mix

Obtaining the rest of the concentrations I just had to alter the volumes from steps c and d, the rest of the steps are exactly the same. The table below shows the volumes used.

Concentration (mol cm³)

3

2.5

2

1.5

1

Volume of acid (cm³)

120

120

120

120

40

Volume of water (cm³)

0

100

80

120

120

Method:

Use ruler and scissors to cut up 3cm magnesium ribbons

Pour some acid into a beaker

Carefully measure out 20cm³ of acid using a measuring cylinder

Place the 20cm³ of acid into the 100ml conical flask

Prepare stopwatch

At the same time place the magnesium ribbon into the flask, seal the flask using the bung connected to the syringe and start the stopwatch.

Take readings of gas produced in the gas syringe every 5 second intervals till no more gas is produced and all the magnesium has reacted

I used a water tank to vary the temperature of the environment. This variable may change the activation enthalpy and rate of reaction between the magnesium ribbon and the acid.

I have used a large (100cm³) gas syringe to allow plenty of space for the gas to collect.

(Above) Figure 1 the apparatus I used to measure the rate of the hydrogen gas produced using a gas syringe

(results)

Results for reactions between magnesium and hydrochloric acid

Using Hydrochloric acid at room temperature:



I was able to calculate the rate of each reaction for each concentration. I did this by taking the initial gradient of each of all 3 tests and finding out the average gradient. The initial gradient is taken from when $t = 0$ (time = 0).

The initial gradient equals the initial rate.

Calculations:

Using 3M HCl:

Time intervals (s)

H₂ gas produced test 1

H₂ gas produced test 2

H₂ gas produced test 3

5

20

18

20

10

35

28

40

15

45

32

41

20

47

38

42

25

53

39

42

30

55

39

42

35

55

39

42

40

55

39

42

45

55

39

42

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$25 \div 5 = 5 \text{ cm}^3 \text{ s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$17.5 \div 5 = 3.5 \text{ cm}^3 \text{ s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$20 \div 5 = 4 \text{ cm}^3 \text{ s}^{-1}$$

Interpretation of all 3 graphs:

The graphs are steep at first, their gradient give the rate of reaction. The steeper the gradient, the faster the reaction. All 3 reactions are at their fastest points at the start, when the concentration of hydrochloric acid is high, before it gets used up. The graphs progressively flatten out, the reason for this is that the hydrochloric acid gets used up and the magnesium ribbon disappears. Once the line becomes horizontal this suggests that the reaction has come to a stop.

Concluding the 3M HCl reactions:

I have calculated the average gradient, thus the average rate of reaction:

Test no. 1 f 4 $\text{cm}^3 \text{ s}^{-1}$

Test no. 2 f 3.5 $\text{cm}^3 \text{ s}^{-1}$ Therefore the average is 4.2 $\text{cm}^3 \text{ s}^{-1}$

Test no. 3 f 4 $\text{cm}^3 \text{ s}^{-1}$

Using 2.5M HCl

Time intervals (s)

H₂ gas produced test 1

H₂ gas produced test 2

H₂ gas produced test 3

0

0

0

0

5

20

19

21

10

32

25

27

15

47

30

37

20

48

35

38

25

52

40

42

30

52

46

47

35

52

46

52

40

52

46

52

45

52

46

52

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$20 \div 5 = 4 \text{ cm}^3 \text{ s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$19 \div 5 = 3.8 \text{ cm}^3 \text{ s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$21 \div 5 = 4.2 \text{ cm}^3 \text{ s}^{-1}$$

Interpretation of all 3 graphs:

Concluding the 2.5M HCl reactions:

I have calculated the average gradient, thus the average rate of reaction:

Test no. 1 f $4 \text{ cm}^3 \text{ s}^{-1}$

Test no. 2 f $3.8 \text{ cm}^3 \text{ s}^{-1}$ Therefore the average is $4 \text{ cm}^3 \text{ s}^{-1}$

Test no. 3 f $4.2 \text{ cm}^3 \text{ s}^{-1}$

The initial steepness of the graphs doesn't last as long as the graphs of the 3M, this is expected at the rate of reaction would be slower due to lower concentration. The graph becomes steadier sooner and stabilizes at an earlier time but longer duration.

Using 2M HCl

Time intervals (s)

H₂ gas produced test 1

H₂ gas produced test 2

H₂ gas produced test 3

5

11

10

10

10

25

19

19

15

27

23

28

20

35

29

35

25

38

38

38

30

42

38

38

35

42

38

38

40

42

38

38

45

42

38

38

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis ÷ change in x-axis

=

$$11 \div 5 = 2.2 \text{ cm}^3 \text{ s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis ÷ change in x-axis

=

$$10 \div 5 = 2 \text{ cm}^3 \text{ s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis ÷ change in x-axis

=

$$10 \div 5 = 2 \text{ cm}^3 \text{ s}^{-1}$$

Interpretation of all 3 graphs:

Again as expected, all graphs took less time to stabilize. However, unusually graph for test 3 took longer time (steepness of the line remained the same for 15seconds) till it started to flatten. This may be a systematic error.

Concluding the 2M HCl reactions:

I have calculated the average gradient, thus the average rate of reaction:

Test no. 1 $2.2 \text{ cm}^3 \text{ s}^{-1}$

Test no. 2 $2 \text{ cm}^3 \text{ s}^{-1}$ Therefore the average is $2.1 \text{ cm}^3 \text{ s}^{-1}$

Test no. 3 $2 \text{ cm}^3 \text{ s}^{-1}$

Using 1.5M HCl

Time intervals (s)

H₂ gas produced test 1

H₂ gas produced test 2

H₂ gas produced test 3

5

9

9

6

10

16

11

10

15

30

16

15

20

35

20

20

25

40

25

23

30

43

30

29

35

47

35

30

40

50

36

32

45

51

37

35

50

52

38

37

55

52

38

38

60

52

38

39

65

52

38

39

70

52

38

39

75

52

38

39

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis ÷ change in x-axis

=

$$9 \div 5 = 1.8 \text{ cm}^3 \text{ s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis ÷ change in x-axis

=

$$9 \div 5 = 1.8 \text{ cm}^3 \text{ s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$6 \div 5 = 1.2 \text{ cm}^3 \text{ s}^{-1}$$

Interpretation of all 3 graphs:

The rate of reaction obviously has decreased significantly, the reaction takes up to 75 seconds to finish and this all supports my hypothesis. Steepness of all 3 gradients of all 3 graphs has decreased and the duration to which the line flattens has also decreased.

Concluding the 1.5M HCl reactions:

I have calculated the average gradient, thus the average rate of reaction:

Test no. 1 f $1.8 \text{ cm}^3 \text{ s}^{-1}$

Test no. 2 f $1.8 \text{ cm}^3 \text{ s}^{-1}$ Therefore the average is $1.6 \text{ cm}^3 \text{ s}^{-1}$

Test no. 3 f $1.2 \text{ cm}^3 \text{ s}^{-1}$

5) Using 1M HCl

Time intervals (s)

H₂ gas produced test 1

H₂ gas produced test 2

H₂ gas produced test 3

10

6

5

4

20

10

9

9

30

16

16

16

40

21

20

19

50

26

24

24

60

28

27

28

70

35

33

32

80

39

34

36

90

42

39

38

100

46

43

42

110

48

45

45

120

52

50

47

130

55

53

50

140

56

55

54

150

57

57

56

160

58

59

58

170

59

59

59

180

59

59

59

190

59

59

59

200

59

59

59

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis ÷ change in x-axis

=

$$6 \div 10 = 0.6 \text{ cm}^3\text{s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis ÷ change in x-axis

=

$$5 \div 10 = 0.5 \text{ cm}^3\text{s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis ÷ change in x-axis

=

$$4 \div 10 = 0.4 \text{ cm}^3\text{s}^{-1}$$

Interpretation of all three graphs:

Concluding the 1M HCl reactions:

I have calculated the average gradient, thus the average rate of reaction:

Test no. 1 f 0.6 cm³s⁻¹

Test no. 2 f 0.5 cm³s⁻¹ Therefore the average is 0.5 cm³s⁻¹

Test no. 3 f 0.4 cm³s⁻¹

Readings were taking at 10second intervals as the reaction time-rate was too slow for 5second intervals

The reaction took 200seconds to complete. This all supports my hypothesis and suggest that the rate of reaction is decreases with the decrease of concentration. The graphs are all less steep and take longer to flatten.

-

From all these calculations of varying concentration of hydrochloric acid reacting with magnesium at room temperature I can create a rate-concentration graph to identify the order of reaction.

Rate

On the left is a table of rate to concentration from the results I have calculated above and in previous pages. Concentrations

0

0

0.5

1

1.6

1.5

2.1

2

4

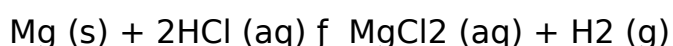
2.5

4.2

3

This graph concludes that the reaction was a first order reaction as it presents a linear line

Using Hydrochloric acid at 70°C



To increase the atmosphere and reaction temperature I placed the conical flask with the 20 cm³ of hydrochloric acid into a water bath and set it to 70 degrees. Using 2 thermometers (one inside the conical flask and the other on the water) I was able to pinpoint the exact point which I could carry out the experiment. Once the temperature in the conical flask, thus the acid, reached 70°C I turned off the electricity so the temperature wouldn't increase and started the experiment. The following are my results and interpretations:

Using 3M HCl at 70°C

Time intervals (s)

H₂ gas produced test no. 1

H₂ gas produced test no. 2

H₂ gas produced test no. 3

5

22

30

25

10

35

40

35

15

42

42

42

20

42

42

42

25

42

42

42

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$22 \div 5 = 4.4 \text{ cm}^3 \text{ s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$30 \div 5 = 6 \text{ cm}^3 \text{ s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis ÷ change in x-axis

=

$$\mathbf{25 \div 5 = 5\text{cm}^3\text{s}^{-1}}$$

Concluding the 3M HCl at 70°C reactions:

I have calculated the average gradient, thus the average rate of reaction:

Test no. 1 f 4.4 cm³s⁻¹ (These gradients are steeper than the room temperature gradients.)

Test no. 2 f 6 cm³s⁻¹ Therefore the average rate is 5.1 cm³s⁻¹

Test no. 3 f 5 cm³s⁻¹

This 3M HCl at 70°C has a higher rate than the 3M HCl reaction at room temperature.

Using 2.5M HCl at 70 °C

Time intervals (s)

H₂ gas produced test no. 1

H₂ gas produced test no. 2

H₂ gas produced test no. 3

5

19

23

21

10

32

35

31

15

41

40

40

20

45

43

43

25

45

43

43

30

45

43

43

35

45

43

43

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis ÷ change in x-axis

=

$$19 \div 5 = 3.8 \text{ cm}^3\text{s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis ÷ change in x-axis

=

$$23 \div 5 = 4.6 \text{ cm}^3\text{s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis ÷ change in x-axis

=

$$21 \div 5 = 4.2 \text{ cm}^3\text{s}^{-1}$$

Concluding the 2.5M HCl at 70°C reactions:

I have calculated the average gradient, thus the average rate of reaction:

Test no. 1 f 3.8 cm³s⁻¹ (These gradients are steeper than the room temperature gradients.)

Test no. 2 f 4.6 cm³s⁻¹ Therefore the average rate is 4.3 cm³s⁻¹

Test no. 3 f 4.2 cm³s⁻¹

This 2.5M HCl at 70°C has a higher rate than the 2.5M HCl reaction at room temperature.

Using 2M HCl at 70°C

Time intervals (s)

H₂ gas produced test no. 1

H₂ gas produced test no. 2

H₂ gas produced test no. 3

5

16

16

15

10

30

29

30

15

40

36

39

20

45

42

44

25

45

44

44

30

45

44

44

35

45

44

44

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$16 \div 5 = 3.2 \text{ cm}^3 \text{ s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$16 \div 5 = 3.2 \text{ cm}^3 \text{ s}^{-1}$$

cm^3

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$15 \div 5 = 3 \text{ cm}^3 \text{ s}^{-1}$$

Concluding the 2M HCl at 70°C reactions:

I have calculated the average gradient, thus the average rate of reaction:

Test no. 1 f 3. 2 cm³s⁻¹ (These gradients are steeper than the room temperature gradients.)

Test no. 2 f 3. 2 cm³s⁻¹ Therefore the average rate is 3. 1 cm³s⁻¹

Test no. 3 f 3 cm³s⁻¹

This 2M HCl at 70°C has a higher rate than the 2M HCl reaction at room temperature.

Using 1. 5M HCl at 70°C

Time intervals (s)

H₂ gas produced test no. 1

H₂ gas produced test no. 2

H₂ gas produced test no. 3

5

13

12

11

10

23

21

21

15

29

27

29

20

34

35

36

25

39

39

41

30

42

42

43

35

42

42

43

40

43

42

43

45

43

42

43

50

43

42

43

Cm3

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$13 \div 5 = 2.6 \text{ cm}^3 \text{ s}^{-1}$$

Initial gradient = initial rate

Therefore, the gradient =

Change in y-axis \div change in x-axis

=

$$12 \div 5 = 2.4 \text{ cm}^3 \text{ s}^{-1}$$

Concluding the 1.5M HCl at 70°C reactions:

I have calculated the average gradient, thus the average rate of reaction:

Test no. 1f 2.6 cm³s⁻¹ (These gradien