

Catalysis of the Zinc-Acid Reaction

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April 21, 2015

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Chapter 1

Abstract

The purpose of this research was to find out information about the zinc and sulfuric acid reaction. I carried out several experiments to find out the order of the reactants and also to find out how effective different catalysts are at increasing the rate of this reaction. I found out that when a catalyst isn't present the sulfuric acid is second order, and when a catalyst is present, both the catalyst and acid are first order, making a second order reaction overall. This experiment was carried out in my college chemistry Laboratory at Long Road Sixth Form College. Inside this coursework, the document also includes in depth chemical ideas about relating topics as well as different ways to carry out the experiment I did.

Chapter 2

Planning

2.1 Chemical Ideas

2.1.1 Rate of Reactions

The definition of a rate of reaction is the rate at which reactants are converted into products. This has the mathematical equation of $Rate = \frac{\Delta Concentration}{Time}$. Finding the rate of reaction is obtained by studying the chemical kinetics of a specific reaction. The rate of reaction gives you important information such as, whether a reaction takes place at all, or how fast a reaction will occur. These chemical ideas are very important in the chemical industry as they allow for companies to reduce the cost and time in producing a product. The balanced chemical reaction only describes the stoichiometry between the reactants present and the amount of products that can be formed from them. Therefore rates are commonly used along side the balanced equation to get a further understanding of the reactions' process.

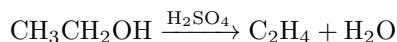
2.1.2 Factors that Affect the Rate of Reaction

Increasing the concentration of the reactants increases the rate of reaction, up to a terminal concentration. This can be explained as two substances physically are unable to react with each other unless their particles come into contact with each other with enough kinetic energy (above the activation energy level). This means that if there is no contact between the particles, or the particles do not have enough kinetic energy, the rate of reaction will be zero. In contrast to this the higher the concentration of the reactants the more likely the particles are to collide above the activation energy per unit time allowing for a reaction to occur.

High temperatures allow for the particles within a system to have a greater kinetic energy. As the kinetic energy of the particles within a system increases, the particles move faster in the enclosed space, thus causing an increased chance of collision per unit time and also allowing for an increased chance of a collision above the activation energy. Virtually all reactions increase in rate if you increase the temperature. Decreasing the temperature has the opposite affect upon the rate of reaction (the rate decreases). In some systems, multiple reactions can take place, with each reaction happening at different temperatures. An example of this is the conversion of ethanol to diethyl ether, this occurs at around 100°C, this is shown below:



However at 100°C a different product is produced, ethylene¹.



Pressure only affects the rate of reaction if the reaction involves gases. Changing the pressure when a reaction involves only solids or liquids will have no effect on the rate of the reaction. Increasing the pressure of a gas is the same as increasing the gases concentration. If a given volume of gas is squeezed into a smaller volume the concentration is higher as shown by the equation:

$$p = \frac{n}{V} \times RT$$

Where:

- p is pressure.
- V is volume.
- n is the number of moles.
- R is the gas constant.
- T is the temperature in K

Therefore this equation shows that as long as the temperature is constant, the pressure is directly proportional to the concentration. If you double one, you will also double the other².

Radiation is a form of energy and therefore it may speed up the rate of reaction as it provides the particles of the reactants with more energy, therefore allowing for more collisions above the activation energy per unit time. As the intensity of radiation increases, the particles absorb more energy and in turn the rate of reaction increases.

Surface area is important in chemical kinetics as increasing the surface area of a reactant will generally increase the rate of reaction. An example of this is iron, in solid form iron is stable enough to be used as solid blocks in building structures; however, in powder form a pyrotechnic composition can be created to form thermite, which when ignited by heat undergoes a vigorous exothermic

reduction-oxidation reaction³. The reason the rate of reaction increases as the surface area of reactants increases is because there is more area exposed that can be hit by other reactant molecules, thus increasing the collisions per unit time.

When two reactants are in the same fluid phase, their particles collide more frequently than when one or both reactants are solids, or when two fluids don't mix. This means that if reactants are in two different phases, collisions between the reactants only occur at the interfaces between the phases. This means that the surface area available for particles to collide is extremely small in comparison to a single homogeneous solution and as discussed above, a smaller surface area equates to a slower reaction time due to the reduced collisions per unit time.

Having a catalyst present in the reaction increases the rate of reaction in both the forward and reverse reactions by providing an alternative pathway of lower activation energy. This means that particles do not need as much kinetic energy to successfully collide together to create a reaction, thus allowing for more successful collisions per unit time. The catalyst, itself remains chemically unchanged. In industry most chemicals are produced using catalysts as they reduce the costs of production. For example, an iron catalyst is used for the Haber process (the process of making ammonia from nitrogen gas and hydrogen gas). This catalyst allows for a higher yield of ammonia without having to increase the temperature or pressure, which would cost a lot of money⁴.

Ion inhibitors such as zinc activated channel antagonists or catalyst poisons decrease the rate of reaction when introduced into a reaction mixture. This is achieved by the ion inhibitors combining with the reactant molecule, thus reducing the number of colliding reactants per unit time, as the surface area of the reactant is covered up. This decrease of successful collisions decreases the rate of reaction. A catalyst poison also reduces the effectiveness of a catalyst. When a reaction occurs with a solid catalyst, a catalyst poison decreases the rate of reaction by accumulating on the surface of the solid catalyst. A common example of a catalyst poison is oxygen and water on an iron catalyst, commonly used in ammonia synthesis, as discussed above. Therefore as there is less surface area on the catalyst, there are less successful collisions per unit time, thus decreasing the rate of reaction.

As a general rule of thumb, introducing low amounts of solvents will increase the concentration, therefore increasing the rate of reaction. Most organic reactions are done in solution for this reason. There are two types of solvents available to use in solution:

- Protic solvents - Contains hydrogen atoms which are free to form hydrogen bonds.
- Aprotic solvents - Contains hydrogen which is all bounded to carbon atoms.

Solvents have an impact on the rate of reaction if they have high dielectric

constants, whereas the non-polar solvents have little or no impact on the rate of reaction. The dielectric constant is a measure of the ability to increase the capacitance of a condenser. These constants increase with the movement of intermolecular dipoles as they align with the external electric field. This means that with a high dielectric constant, the ability of the molecule to spread the charges increases. This is a macroscopic property and influences how the solvent interacts with the solutes. This then allows for protic solvents to increase the rate of reaction along with other solvent properties such as: Empirical polarity criteria, solvent electrophilicity, solvent nucleophilicity and solvent cohesivity⁵.

More about enthalpy level diagrams and Maxwell-Boltzmann distribution can be found in section 2.1.12 on page 20. This outlines the specific affect that these factors have on the mechanism behind speeding up the rate of reactions.

2.1.3 Rate Equations

The rate equation shows you how fast a reaction occurs with respect to the reactants/products. The general rate equation can be expressed as:

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

Where:

- 'k' is the rate constant
- '[A]' is the concentration of A.
- '[B]' is the concentration of B.
- 'm' and 'n' are the orders of A and B, respectively.

The values of m and n can only be determined experimentally and do not directly relate to the moles of the substance. The addition of m and n give you the overall order of the reaction. As the rate equation is a differential equation, it can be integrated to give the integrated rate equation. This links the concentrations of individual reactants or products over time.

The rate constant can be found out from the Arrhenius equation, this can be expressed as:

$$k = Ae^{-\frac{E_a}{RT}}$$

Where:

- 'k' is the rate constant.
- 'A' is the frequency factor.
- 'e' is the mathematical quantity (2.71828).
- ' E_a ' is the activation energy.
- 'R' is the gas constant.

- 'T' is the temperature in kelvin.

From this equation you can show that changing the condition of the reaction also changes the value of the rate constant. For example, if the temperature increases, so does the rate constant. By assuming that the activation energy is 50000 J mol⁻¹ and the value of the gas constant is 8.31 J K⁻¹ mol⁻¹, at 20°C (293 K) the value of the fraction of the Arrhenius equation would be:

$$e^{-\frac{E_a}{RT}} = e^{-\frac{50000}{8.31 \times 293}}$$

$$e^{-\frac{E_a}{RT}} = 1.21 \times 10^{-9}$$

By raising the temperature by 10°C to 30° (303 K) the value of the fraction would change to:

$$e^{-\frac{E_a}{RT}} = e^{-\frac{50000}{8.31 \times 303}}$$

$$e^{-\frac{E_a}{RT}} = 2.38 \times 10^{-9}$$

By changing the temperature of the reaction by just 10°C the rate of reaction has almost doubled (1.21 × 10⁻⁹ to 2.38 × 10⁻⁹).

As discussed previously in section 2.1.2 on page 5 the catalyst provides a route for the reaction of lower activation energy. By changing the 20°C reactions activation energy to 25000 J mol⁻¹ the equation would be:

$$e^{-\frac{E_a}{RT}} = e^{-\frac{25000}{8.31 \times 293}}$$

$$e^{-\frac{E_a}{RT}} = 3.47 \times 10^{-5}$$

From this equation you can see that reducing the activation energy required for a reaction to occur, the number of molecules that react, significantly increase¹⁹.

2.1.4 Orders of Reactions

Zero-Order Reactions

In a zero-order reaction the rate appears to be completely independent of the concentration of the reactants. This means that increasing or decreasing the concentration of the reactants will not have an impact on the rate of the reaction. This leaves the rate equation to be:

$$\text{Rate} = k$$

In respect to the zero-order reactant; however the reactants that have an impact on the rate of reaction are still shown in the equation. For example, if [A] doesn't change the rate of reaction when its concentration is varied but [B] does (as a first order reactant), the equation would be:

$$\text{Rate} = k[B]$$

This equation shows that [A] has no impact on the rate of reaction, but [B] is a first-order reactant, so overall the rate equation would be first-order.

This equation can also be represented as a rate vs concentration graph. This is shown below and illustrates that the rate of reaction does not vary when the concentration of reactants increase.

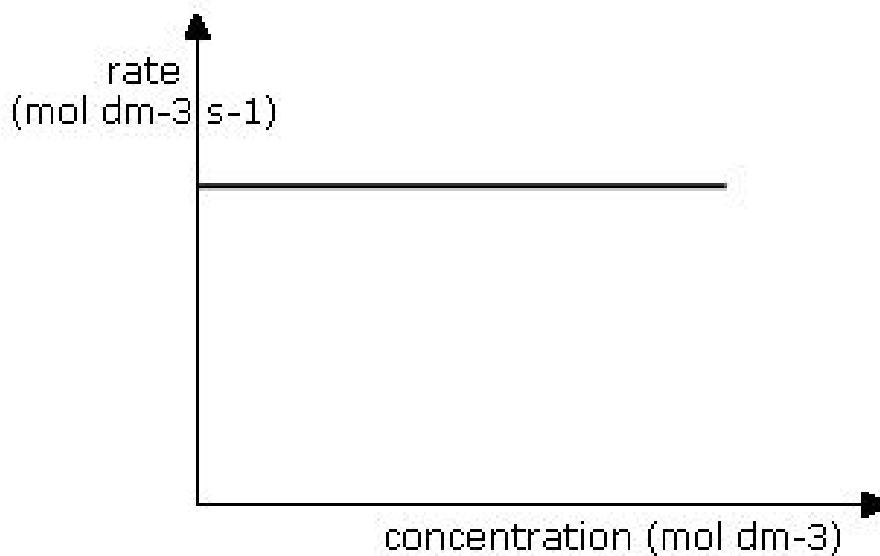
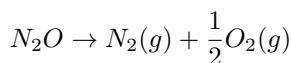


Figure 2.1: Rate vs Concentration Zero-Order Graph

From the equation you can see that the rate only depends on the rate constant of the reaction. Zero-order reactions occur when the reactant has become exhausted; however before this point is reached, the reaction will follow a different rate law instead of falling to zero directly. There are two general conditions of a reaction that can result in a zero-order reaction:

- Only a small amount of the reactant molecule are able to react e.g small reacting area/wrong state.
- When two or more reactants are involved in the reaction the concentrations of some reactants are greater, and therefore overwhelm the other reactants.

This usually occurs when a reaction is catalysed by heterogeneous catalysis or to an enzymes specific active site. An example commonly exhibited of this is the decomposition of nitrous oxide.



The reaction on the previous page occurs in the presence of a hot platinum wire catalyst. In this reaction the N_2O molecules are limited to the surface area available on the catalyst. Once all of the surface area is occupied at a given time the reaction has an order of zero. This continues until a part of the surface area is freed up after the products desorb from the catalyst surface and diffuse away⁶. Zero-order reactants have no units as they do not impact the reaction.

First-Order Reactions

A first-order reactions' rate depends linearly on one single reactants concentration; however this does not mean another reactant couldn't have an impact on the rate. For a single first-order reactant can be shown by the differential rate equation:

$$\text{Rate} = k[A]$$

The rate constants unit can be obtained mathematically by rearranging the equations units to equal 'k'. When this is done the rate constants unit is s^{-1} . This equation can also be represented as a rate vs concentration graph. This is shown below and illustrates that the rate of reaction varies only on one single reactants concentration.

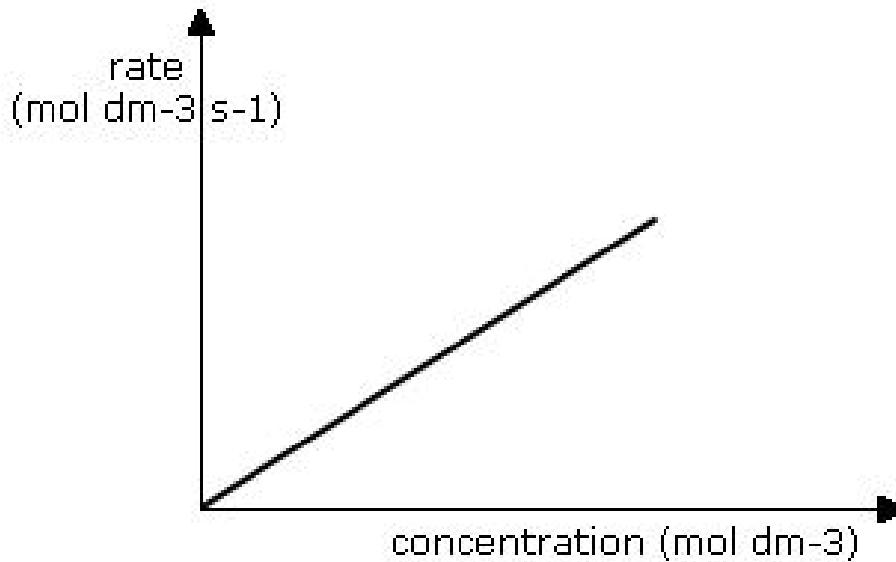


Figure 2.2: Rate vs Concentration First-Order Graph

Second-Order Reactions

In a second order reaction, the rate equation can be displayed in two ways depending on how the reaction is carried out.

1. $\text{Rate} = k[A][B]$
2. $\text{Rate} = k[A]^2$

Both of these rate equations exhibit a second-order reaction. The first example occurs when the reactants are different and combine in a single step. The first equation shows that if you double A then the rate of reaction is doubled, and the same happens if you double B. For the second example the reactants would have to be identical, and combine in a single step. Therefore doubling A would quadruple the rate of reaction. This means that the rate vs concentration graph would appear to look like the graph below.

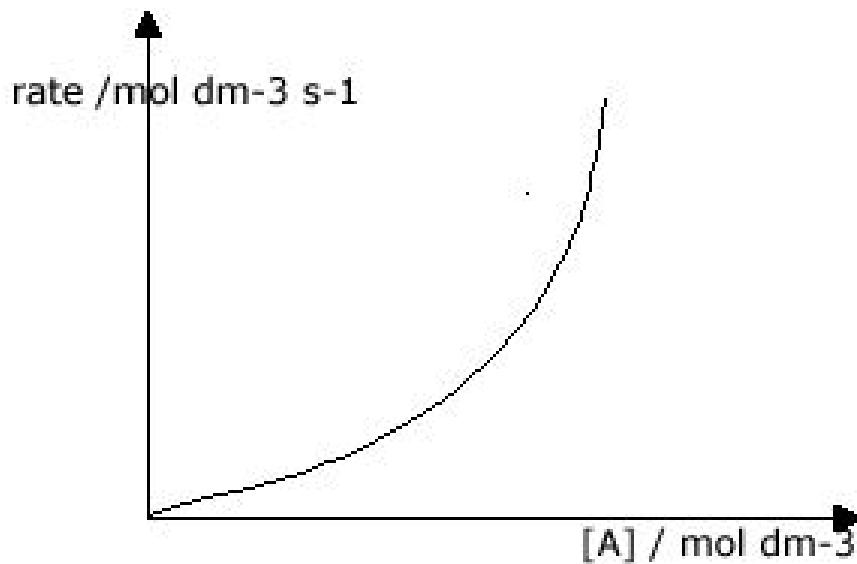


Figure 2.3: Rate vs Concentration Second-Order Graph

The rate constant's unit can be obtained mathematically by rearranging the equations units to equal 'k'. When this is done the rate constant's unit is $\text{mol dm}^{-3}\text{s}^{-1}$.

Third-Order Reactions

Some reactions have multiple reactants, or reactants that are second-order and first-order, thus resulting in a third-order reaction overall. Examples of this include, but aren't limited to:

$$\text{Rate} = k[A][B]^2$$

$$\text{Rate} = k[A][B][C]$$

The unit of k for a third-order reaction is $\text{mol}^{-2} \text{ dm}^{-6} \text{ s}^{-1}$.

2.1.5 Finding Rates

Rates have to be determined experimentally. The average rate of reaction is obtained by taking the change in concentration over a specific time period. This value is only an approximation of the rate of reaction in that specific time period. This means that this method does not give you the specific rate values throughout that time interval and in some cases, that rate value is not even equal at any time throughout that time interval.

Another method called the instantaneous rate of reaction, in contrast to the average rate of reaction allows for a more accurate value to be obtained. This method is defined as the change in concentration of an infinitely small time interval. The derivative mathematical equation for this is $\frac{d[\text{Concentration}]}{dT\text{ime}}$. By graphing the concentration of the experiment over time you are then able to draw a tangent line at a specific time value. The gradient of the tangent is then directly proportionate to the rate of reaction. You can also mix these two methods to get an average rate which is closer to the value of the instantaneous rate. This is done by measuring the change in concentration over a very small time period, two or more times to obtain an average. Again the rate of reaction for that time period is determined by the slope of the tangent lines⁷.

Justification of Chosen Method

For my experiment I will be finding the rate of reaction using the following steps:

1. Plot a graph of the volume of hydrogen against time.
2. From the graph draw a tangent to the line at the initial point.
3. Calculate the gradient of the tangent by using the equation:
4. The gradient is equal to the rate of reaction.

I have chosen this method as it is the method which will give me the most accurate values for the specific point in time which I wish to measure.

2.1.6 pH

The pH scale is composed of two extremes that describes a chemical property about the substance being tested, these extremes are called acids and bases. Mixing acids and bases together will induce a neutralisation reaction which can cancel out their extreme effects. A substance which is neither acidic nor basic is called a neutral substance. The pH scale ranges from 0 to 14, with 0 being as acidic as possible and 14 being as basic as possible. Neutral has a corresponding pH of 7 and therefore anything below 7 is acidic and anything above 7 is basic. The pH scale is illustrated by the following chart.

Concentration of hydrogen ions compared to distilled water	pH	Examples of solutions at this pH
10,000,000	pH = 0	battery acid, strong hydrofluoric acid
1,000,000	pH = 1	hydrochloric acid secreted by stomach lining
100,000	pH = 2	lemon juice, gastric acid, vinegar
10,000	pH = 3	grapefruit, orange juice, soda
1,000	pH = 4	tomato juice, acid rain
100	pH = 5	soft drinking water, black coffee
10	pH = 6	urine, saliva
1	pH = 7	"pure" water
1/10	pH = 8	sea water
1/100	pH = 9	baking soda
1/1,000	pH = 10	Great Salt Lake, milk of magnesia
1/10,000	pH = 11	ammonia solution
1/100,000	pH = 12	soapy water
1/1,000,000	pH = 13	bleaches, oven cleaner
1/10,000,000	pH = 14	liquid drain cleaner

Figure 2.4: The pH scale

The pH scale is a man made scale which is used to measure the concentration of hydrogen ions, each concentration is given a corresponding place on the scale (pH). pH is mathematically defined as the negative logarithm of the hydrogen ion concentration. As a result of this we can determine that the pH scale is

logarithmic, therefore each value above/below the neutral value (7) is ten times more basic/acidic respectively. For example pH 6 is ten times as acidic as pH 7 and pH 5 is one hundred times as acidic than pH 7. The mathematical equation for working out pH is illustrated below.

$$\bullet \quad pH = -\log[H^+]$$

There are many indicators used to find out the pH of substances. A table of common indicators with their properties are displayed in the following table.

Indicator	Low pH color	Transition pH range	High pH color
Thymol blue (first transition)	Red	1.2 – 2.8	Yellow
Methyl red	Red	4.4 – 6.2	Yellow
Bromothymol blue	Yellow	6.0 – 7.6	Blue
Thymol blue (second transition)	Yellow	8.0 – 9.6	Blue
Phenolphthalein	Colorless	8.3 – 10.0	Fuchsia

Figure 2.5: List of pH Indicators

Universal indicator contains all of the chemicals above, all mixed into a single solution or into universal indicator paper. This allows for a continuous colour change from about pH 2 to pH 10. Visual comparison of the colour of the universal indicator and a standard colour chart give a rough reading of the pH of the substance being measured, usually to the nearest whole number.

2.1.7 Acids

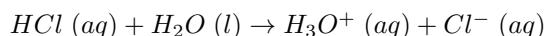
The Brønsted–Lowry theory defines an acid as a proton H⁺ donor. They have specific chemical and physical properties. Some of these are detailed in the table overleaf.

Acid Properties
Piercing pain when placed on a wound.
Tastes sour.
Colourless when placed in phenolphthalein.
Turns blue litmus paper red.
Has a pH value of less than 7.
Produces hydrogen gas when reacted with metals.
Produces carbon dioxide when reacted with carbonates ⁸ .

Acids can be strong or weak depending on their ability to dissociate in water. Strong acids can fully dissociate in water by ionising into H⁺ ions and the additional anion. These strong acids are usually inorganic acids such as: HCl, HNO₃ and H₂SO₄. Weak acids also dissociate, producing H⁺ ions and an anion; however these only partially dissociate. This means that in water some molecules will not dissociate as the reaction between the ions and the acid is reversible and this equilibrium is in favour of the acid. Weak acids are usually organic acids, such as: Ethanoic acid and benzenol.

The strength of an acid is not related to its concentration. Both weak and strong acids can be in high concentrations and low concentrations. The concentration refers to the number of moles of the specific acid in a given solution.

When an acid dissociates in water, an oxonium ion is formed from the hydrogen in the acid and another ion is formed from the rest of the acid molecule. For example, hydrochloric acid reacts with water to form an oxonium ion and a chlorine ion. This is shown below.



2.1.8 Catalysts

Catalysts are used to speed up the rate of a chemical reaction by finding an alternate pathway of lower activation energy. They do not get used up in a chemical reaction and can be recovered chemically unchanged, often only tiny amounts of catalysts are needed for the catalysts optimum potential⁹. Although catalysts are not used up in the reaction, they may be destroyed, deactivated or inhibited by a different process.

A chemical reaction involving a catalyst allows for less energy to be required to reach the transition state, but the total free energy over the process from reactants to products does not change¹⁰. A catalyst can be involved in multiple steps in a chemical reaction. Catalytic reactions occur in the same way as normal chemical reactions when looking at the kinetics of the reaction (the rate depends on successful collisions of particles). Catalysts do not change the position of the equilibrium. This means, although the initial yield seems to be greater, over time the overall yield would be the same with or without a catalyst.

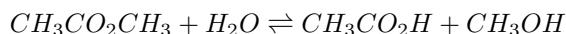
Heterogeneous catalysis occurs when the catalyst acts in a different phase to that of the reactants. Generally speaking most heterogeneous catalysts are solids that speed up a reaction involving liquids or gases. A heterogeneous catalyst has active sites which allows a reaction to occur between the reactants. This happens over a 4 step mechanism.

1. Reactants adsorb onto the catalysts surface.
2. Bonds within the reactants molecules weaken and break.
3. New bonds form to make products.
4. The product desorbs from the catalysts surface and diffuses away.

The active site of the catalyst is only in specific areas due to imperfections in the structure of the catalysts surface at an atomic level. It has been shown that an irregular catalytic surface appears to be more effective than a flat plane surface¹¹

Homogeneous catalysis is a chemical reaction that involves a catalyst that is in the same phase of the reactants. Normally homogeneous catalysts are dissolved in a solvent with the reactants. Once the reaction is complete the catalyst is usually released in their initial form, this is why the catalyst is not present in the net reaction equation.

Homogeneous acid catalysis is one of the most common forms of homogeneous catalysis. This is because water is an extremely common solvent which happens to partially dissociate to produce H⁺ ions. An example of this catalysis is the hydrolysis of esters. As acids catalyse the hydrolysis of esters, carrying out the reaction in water will naturally speed up the reaction, as without any catalytic action the hydrolysis of esters takes a long time. An example acid catalysis is shown below.



2.1.9 Transition Metal Catalysts

Transition metals are the group of metals in the d block of the periodic table. They are defined as an atom in the d block of the periodic table that has a partially filled d sub-shell and that forms a stable ion. They are divided into three groups:

- First row transition metals.
- Second row transition metals.
- Third row transition metals.

Many of these transition metals have catalytic properties, either as a solid metal itself, or a mixed compound. These catalysts work in the same way and are defined in the same way as discussed in the 'Catalyst' section previously.

Transition metals have a number of physical and chemical properties, these are listed in the table below.

Physical Properties	Chemical Properties
High melting point	Forms compounds with a variety of oxidation states
High boiling point	Catalytic activity
Good thermal conductivity	Strong tendency to form complexes
Good electrical conductivity	Formation of coloured compounds
High density	
Hard/strong, especially in alloys	

In my experiment I am using copper sulfate as the catalyst to speed up the zinc and sulfuric acid reaction. The reaction of zinc and sulfuric acid at room temperature is rather slow in comparison to the zinc and hydrochloric acid reaction. Hence why I am using the slower reaction to study the catalytic affect on the reaction. I know that the copper is the catalytic component of copper sulfate as the reaction $\text{Zn} + \text{CuSO}_4 \rightarrow \text{Cu} + \text{ZnSO}_4$ shows this by being the single remaining component. This works as zinc and copper present in sulfuric acid creates an electrolytic cell. In this electrolytic cell the zinc acts as the anode as it is more reactive and the copper acts as the cathode, as it is not as reactive as zinc¹². Therefore hydrogen ions are going to be reduced to form hydrogen gas (one of the products of a metal and acid reaction) at the copper electrode. This leads to an increased production of hydrogen gas. Without the copper acting as a catalyst (by forming an electrolytic cell) the hydrogen would have to be produced by the zinc and the hydrogen ions from the acid at the surface of the zinc. This, for obvious reasons, produces much less hydrogen than if the copper transition metal catalyst was present¹³.

2.1.10 d Orbitals

An orbital is defined as a region in an atom where there is a high probability of finding an electron. In total there are 5 different d orbitals. These orbitals are referred to as: d_z^2 , d_{xy} , d_{xz} , d_{yz} , and $d_{x^2-y^2}$. These d orbitals are the reason that transition metals have specific physical and chemical properties, as discussed above. As transition metals have a partially filled outermost d sub shell they are able to donate and accept electrons easily, which is why they exhibit catalytic activity. Overleaf shows a representation of the five d orbitals.

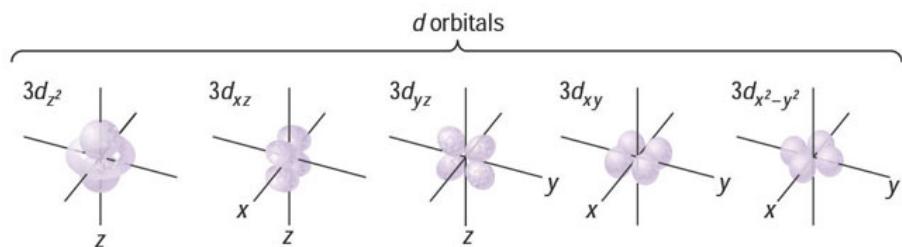


Figure 2.6: d Orbital representations

Copper has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$. But when it forms an ion (Cu^{2+}) it loses its 4s electron and also one of its 3d electrons. The electron configuration is now $1s^2 2s^2 2p^6 3s^2 3p^6 3d^9$. This means that the copper ion fits the definition of a transition metal catalyst. The electron representation of the copper ion is shown below.

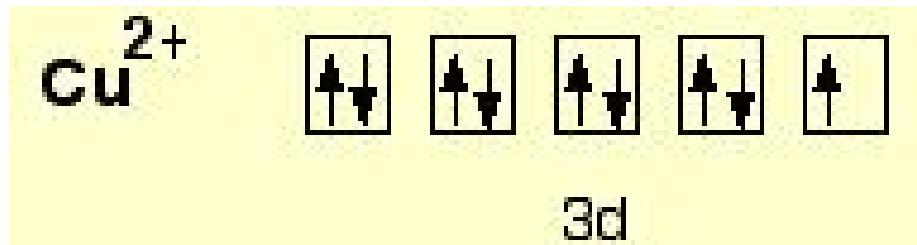


Figure 2.7: Copper Ion D Orbital Electrons

2.1.11 Complexes and their Properties

A complex is a species that consists of a central metal atom or ion that is surrounded by ligands. Complexes are formed when ligands are bonded to a transition metal by a dative covalent bond. The number of d electrons on the central metal atom or ion is important as transition metals can exist in many different oxidation states which therefore means a different amount of d electrons in the outer shells. By oxidising or reducing the central metal atom or ion, with the number and type of ligands unchanged, many of the complexes change. These properties include:

- Reactivity.
- Magnetic Activity.

- Stability.
- Stereochemistry.
- Spectroscopic properties.

Complexes contain a coordination number which identifies the number of dative bonds that are formed from the surrounding ligands to the central metal atom or ion. Usually the coordination number is: 2, 4 or 6. Each coordination number causes a different shaped complex. These different shapes are outlined below.

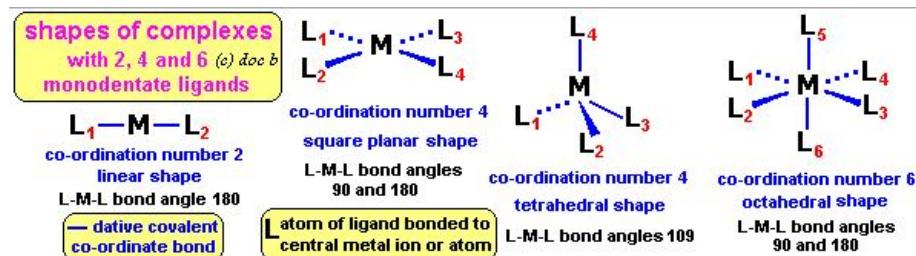


Figure 2.8: Complex Shapes and Structures

2.1.12 Maxwell-Boltzmann Distributions

The Maxwell-Boltzmann distribution is used to identify the kinetic energy distribution for a reaction. The mathematical equation for working out the Maxwell-Boltzman distribution is equivalent to:

$$\frac{dN}{N} = \left(\frac{m}{2\pi k_B T} \right)^{1/2} e^{\frac{-mv^2}{2k_B T}} dv \quad ^{17}$$

Where:

- dN/N is the fraction of molecules moving at velocity v to $v + dv$.
- m is the mass of the molecule.
- k_B is the Boltzmann constant.
- T is the absolute temperature.

As discussed before, the factors that affect the rate of reaction can be illustrated by the Maxwell-Boltzmann distribution graphs. By increasing the temperature the rate of reaction increases as more particles have a kinetic energy that exceeds the activation energy of the reaction. The graph below shows two curves, both of which are at different temperatures (300 K and 310 K). As shown by the graph overleaf the increase in temperature results in the peak of the graph shifting to the right and decreasing in height. This shows that there is an increased

frequency of successful collisions. Increasing the temperature does not change the area of the graph as the number of particles to begin with remain the same.

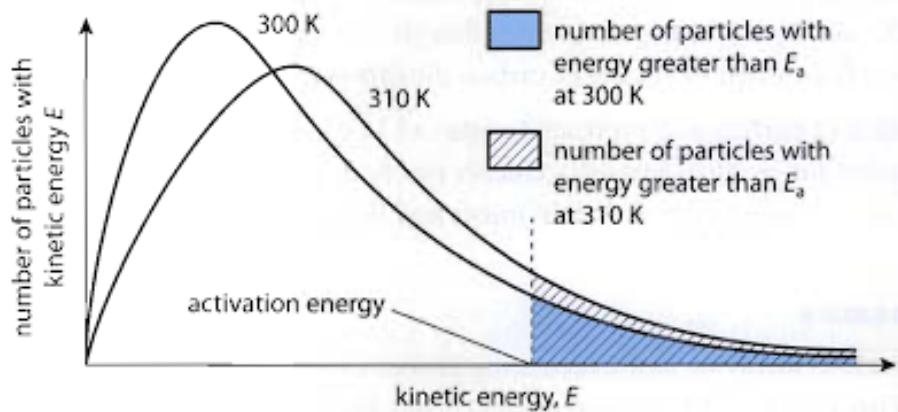


Figure 2.9: Temperature affect on the Maxwell-Boltzmann Distribution Graph

The effect of adding a catalyst provides an alternative route of lower activation energy, thus speeding up the rate of reaction. An in depth analysis of catalyst affect can be found in the catalyst and factors that affect reactions in sections 2.1.8 and 2.1.2 on pages 16 and 5. As shown by the graph overleaf, the lower activation energy leads to more particles having kinetic energy greater than the activation energy. Therefore more particles react, which ultimately leads to a faster reaction.

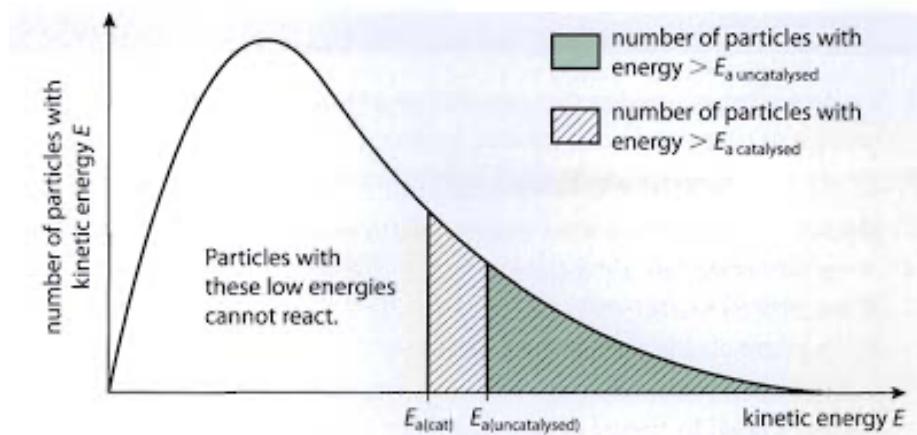


Figure 2.10: Catalyst affect on the Maxwell-Boltzmann Distribution Graph

2.1.13 Enthalpy Level Diagrams

Enthalpy is the heat energy content of a substance. The net heat energy transferred to a system from the surroundings to a system at constant pressure is known as the enthalpy change (ΔH). This can be worked out mathematically by applying the formula:

$$\Delta H^\circ = \sum H_f^\circ \text{products} - \sum H_f^\circ \text{reactants}$$

An exothermic reaction would produce a negative ΔH value as the enthalpy of the reaction system is decreasing. This occurs as heat energy is given out from the system to the surroundings. Therefore the greater the heat energy released from the reaction, the more exothermic the reaction and therefore the further above the reactants enthalpy level would be in comparison to the products.

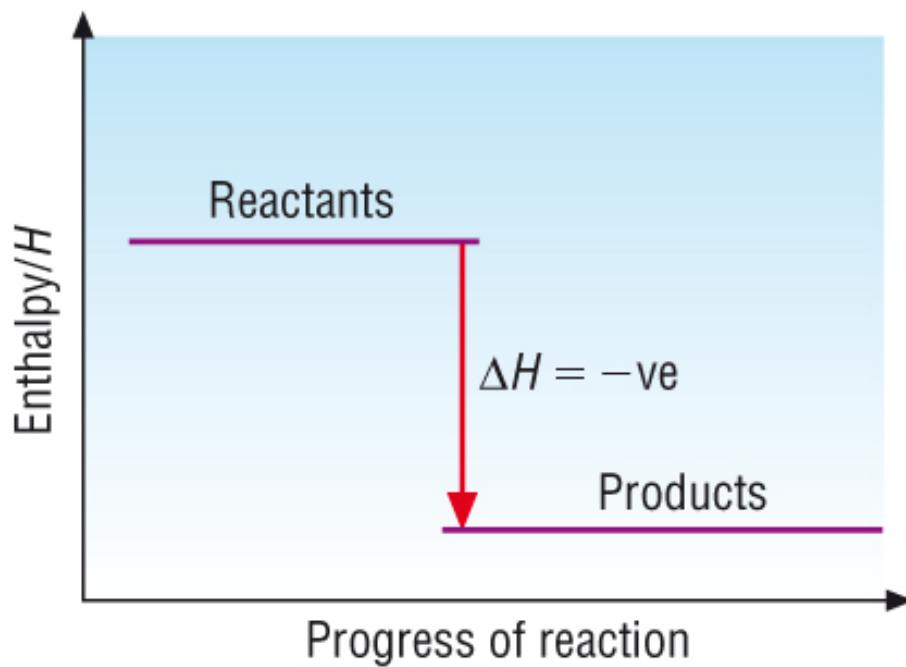


Figure 2.11: Exothermic Enthalpy Diagram

An endothermic reaction would produce a positive ΔH value as the enthalpy of the reaction system is increasing. This occurs as heat energy is absorbed from the surroundings. Therefore the lower the heat energy (the more heat energy that is taken into the reaction), the more endothermic the reaction would be and therefore the further below the reactants enthalpy level would be in comparison to the products.

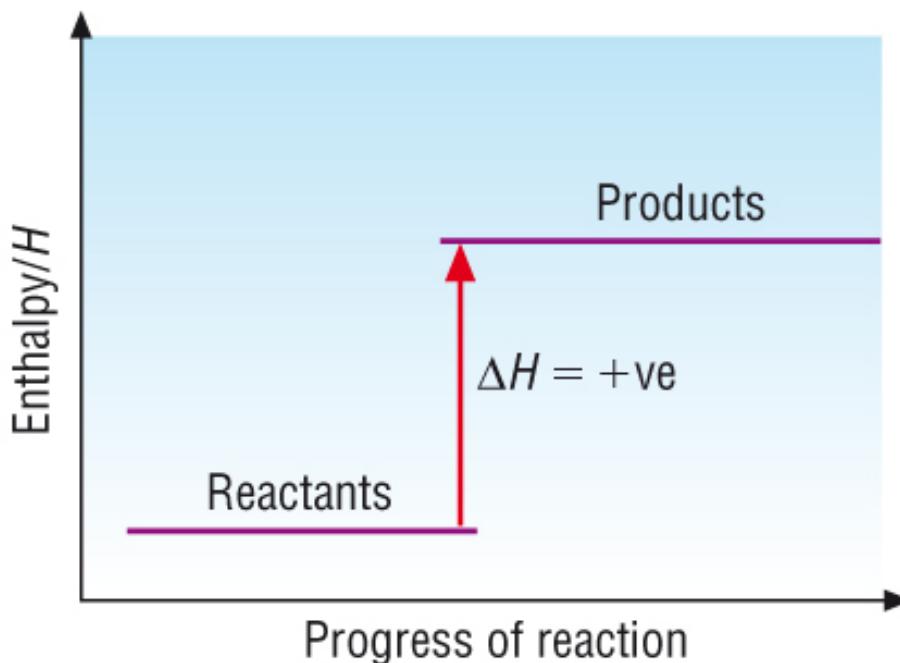


Figure 2.12: Endothermic Enthalpy Diagram

All enthalpy equations are carried out under standard conditions which has 3 criteria.

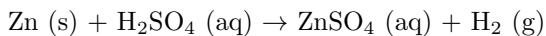
1. 298K.
2. 1ATM.
3. 1 Molar Solutions.

The state symbols also need to be involved as the molecules have different amounts of energy in different states.

Catalysts do not change the enthalpy value of the reaction. Catalysts only change the activation energy required to start a successful collision.

2.1.14 Metal and Acids

Usually when metals and acids are mixed together hydrogen gas is produced along with a salt. With the zinc and sulfuric acid reaction zinc sulfate is the salt produced. The hydrogen produced in a metal and acid reaction can be observed from the effervescing of the solution. The chemical equation for the zinc and sulfuric acid reaction is shown overleaf.



2.1.15 Mechanisms

Reaction mechanisms are used to identify what is happening during a reaction on a molecular level. An elementary process is an individual step of the reaction mechanism which describes a moment where one or more of the molecules involved in the reaction interact with another molecule. Multiple elementary processes can be collectively combined to show the overall reaction. Elementary processes can occur at a unimolecular, bimolecular or termolecular level (different number of molecules colliding). A unimolecular elementary step would be a first order rate law, whereas a bimolecular elementary step has a second order rate law. And finally a termolecular elementary step has a third order rate law. In a bimolecular elementary process the slow step of the reaction is the rate determining step as it occurs more slowly. When the reaction has enough energy the reaction moves to a faster step which occurs more quickly.

All of the elementary steps must add up to give the overall balanced equation for the reaction, and the rate law for the rate determining step must agree with the experimentally determined rate equation.

In my analysis section I will identify the mechanism for the reactions which I am carrying out.

2.1.16 Oxidation States

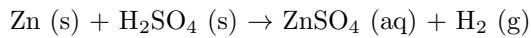
An elements oxidation state represents the number of electrons that an atom can share, gain or lose when bonding with another atom. Oxidation states can be used to determine the changes that occur to an atom when a redox reaction occurs. There are several rules that must be followed in order. If two rules conflict then the rule that comes first holds true.

1. Any individual atom had an oxidation state of 0.
2. The total sum of the oxidation state of all atoms is equal to the net charge of the species.
3. Group 1 metals have an oxidation state of +1 and group 2 metals have an oxidation state of +2.
4. Fluorine has an oxidation state of -1 in a compound.
5. Hydrogen has an oxidation state of +1 in a compound.
6. Oxygen has an oxidation state of -2 in a compound.
7. Chlorine had an oxidation state of -1

There are exceptions to the rules shown above. These are outlined below:

- Hydrogen in metal hydrides has an oxidation state of -1.
- Oxygen in peroxides has an oxidation state of 0.
- Oxygen in F₂O has an oxidation state of +2.
- Chlorine in compounds with fluorine or oxygen has a variety of possible oxidation states.

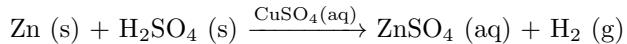
In my experiment I used zinc, sulfuric acid and copper (II) sulfate (VI). My two main experiment series involves the non-catalysed and catalysed reaction between zinc and sulfuric acid. The two equations are shown below, respectively.



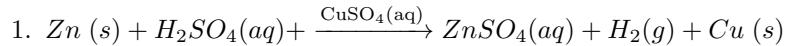
The SO₄ ion is a spectator ion as it doesn't change oxidation state, in the above reaction and the two ionic equations can be represented as:

1. $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$
2. $2\text{H}^+ 2e^- \rightarrow \text{H}_2$

The two ionic equations show that there are two oxidation state changes involved in the uncatalysed reaction between zinc and sulfuric acid. This involves zinc being oxidised to a zinc (II) and hydrogen ions being reduced to hydrogen gas.



Again the SO₄ ion is a spectator ion. The catalysed reaction is a bit more complex as if there were any copper sulfate in the solution, solid copper would be deposited. This is shown below.



There are three oxidation state changes involved in the catalysed reaction, and these are shown below.

1. $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$
2. $2\text{H}^+ 2e^- \rightarrow \text{H}_2$
3. $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$

The three ionic equations show that there are three oxidation state changes involved in the catalysed reaction of zinc and sulfuric acid with copper sulfate. This involves zinc being oxidised to a zinc (II), hydrogen ions being reduced to hydrogen gas and copper (II) being reduced to copper.

2.2 Inventory

2.2.1 Equipment List

- 250 cm^3 conical flask.
- Bung fitted to a glass tube.
- Gas Syringe.
- Thermometer.
- Clamp Stand.
- Range of volumetric flasks.
- Volumetric Pipette.
- Weighing Scales.
- Weighing Boat.

2.2.2 Chemical List

- Distilled Water.
- 5.896g Hydrated Copper Sulfate (s) .
- 4.0 mol dm^{-3} Sulfuric Acid (aq) .
- Granulated Zinc (s) .
- Powdered Zinc (s) .
- Mixture of Different Catalysts.

2.2.3 Preparing Chemicals

Preparing the Copper Sulfate

I am using a hydrated form of copper sulfate ($CuSO_4 \cdot 5H_2O$), this is a solid, bright blue coloured salt. I decided to use the hydrated form so that I could make up different solutions of different concentrations easily. Copper sulfate forms an aqueous solution when water is introduced. This is an exothermic reaction which gives the complex: $[Cu(H_2O)_6]$. This complex has an octahedral molecular geometry¹⁵ and displays paramagnetic properties¹⁶.

To make the desired concentration of copper sulfate from the hydrated salt I have to dissolve it in a specified volume of water. The molecular mass of the hydrated form of copper sulfate is 249.6, therefore I have ordered 5.896 grams to make up a $200cm^3$, 0.1 molar solution of copper sulfate. This gives me enough

copper sulfate to use for my whole experimental series, for each experimental series I made up a new standard solution. To make up the solution I weighed out 5.896g to ensure that the mass was correct and added that to a 200 cm³ volumetric flask and then filled up the volumetric flask with distilled water until the meniscus of the solution reached the line. I will then shake the solution until all of the copper sulfate has been dissolved completely to make up a 200 cm³, 0.1 molar solution.

The table below shows all of the solutions involving copper sulfate that I made up and how much of each chemical I used to make the final solution.

Final Solution	Copper Sulfate Added	Sulfuric Acid Added	Distilled Water Added
200 cm ³ of 0.01 Molar Copper Sulfate and 0.2 Molar Sulfuric Acid	20 cm ³ of 0.1 Molar	10 cm ³ of 4.0 Molar	170 cm ³
200 cm ³ of 0.02 Molar Copper Sulfate and 0.2 Molar Sulfuric Acid	40 cm ³ of 0.1 Molar	10 cm ³ of 4.0 Molar	150 cm ³
200 cm ³ of 0.03 Molar Copper Sulfate and 0.2 Molar Sulfuric Acid	60 cm ³ of 0.1 Molar	10 cm ³ of 4.0 Molar	130 cm ³
200 cm ³ of 0.04 Molar Copper Sulfate and 0.2 Molar Sulfuric Acid	80 cm ³ of 0.1 Molar	10 cm ³ of 4.0 Molar	110 cm ³
200 cm ³ of 0.05 Molar Copper Sulfate and 0.2 Molar Sulfuric Acid	100 cm ³ of 0.1 Molar	10 cm ³ of 4.0 Molar	90 cm ³
200 cm ³ of 0.06 Molar Copper Sulfate and 0.2 Molar Sulfuric Acid	120 cm ³ of 0.1 Molar	10 cm ³ of 4.0 Molar	70 cm ³
200 cm ³ of 0.1 Molar Copper Sulfate and 0.2 Molar Sulfuric Acid	20cm ³ of 1.0 Molar	10 cm ³ of 4.0 Molar	170 cm ³
200 cm ³ of 0.1 Molar Copper Sulfate and 0.4 Molar Sulfuric Acid	20cm ³ of 1.0 Molar	20 cm ³ of 4.0 Molar	160 cm ³
200 cm ³ of 0.1 Molar Copper Sulfate and 0.6 Molar Sulfuric Acid	20cm ³ of 1.0 Molar	30 cm ³ of 4.0 Molar	150 cm ³
200 cm ³ of 0.1 Molar Copper Sulfate and 0.8 Molar Sulfuric Acid	20cm ³ of 1.0 Molar	40 cm ³ of 4.0 Molar	140 cm ³
200 cm ³ of 0.1 Molar Copper Sulfate and 1.0 Molar Sulfuric Acid	20cm ³ of 1.0 Molar	50 cm ³ of 4.0 Molar	130 cm ³
200 cm ³ of 0.1 Molar Copper Sulfate and 1.2 Molar Sulfuric Acid	20cm ³ of 1.0 Molar	60 cm ³ of 4.0 Molar	120 cm ³

Preparing the Sulfuric Acid

For my project I will need to produce a range of standard solutions for sulfuric acid to use for my reactions in order to determine the rate and order of the reactions. To do this I will rearrange the equation $[Final] = [Initial] \times \frac{InitialVolume}{FinalVolume}$ to make the initial volume the subject of the equation, this is shown below.

$$InitialVolume = \frac{[Final] \times FinalVolume}{[Initial]}$$

As I am ordering 4.0 molar sulfuric acid, my initial concentration will be 4.0 molar and I will make up 200 cm³ of the standard solution so that I will be able to carry out 5 repeats which allows me to repeat the experiment twice more than I am planning to, in case of experimental error or anomalies. Therefore my Final Volume will be 200 cm³ and my final concentration will correspond to the concentration which I want to make up the solution to. For example, with my starting concentration of 2.0 molar the equation will look like:

- $InitialVolume = \frac{2.0\text{Molar} \times 200\text{cm}^3}{4.0\text{Molar}}$
- $100\text{cm}^3 = \frac{2.0\text{Molar} \times 200\text{cm}^3}{4.0\text{Molar}}$

This means that my initial volume of 4.0 molar sulfuric acid will be 100 cm³. To make up the solution I will use a volumetric pipette to measure out 100 cm³ of the 4.0 molar sulfuric acid and add that to a 200 cm³ volumetric flask. I will then fill up the volumetric pipette (up to when the meniscus of the solution rests on the line) with distilled water. I will then shake the volumetric flask to mix up the solution and therefore make the solution to 2.0 molar sulfuric acid.

When I make up my sulfuric acid solution with catalysts I will take 10 cm³ of 0.1 molar catalyst and replace the desired amount of the distilled water with the specific catalyst. This will not change the concentration of the acid as the solution is still made up to the same volume. I will use a volumetric pipette to measure the desired volume of catalyst and add it to the volumetric flask after the acid. Again, I will fill up the volumetric flask to the line with distilled water after carefully measuring the desired volume of the catalyst and acid.

The table below outlines all of the sulfuric acid concentrations I made up and how much of each chemical I used to make the final solution.

Final Solution	Sulfuric Acid Added	Distilled Water Added
200 cm ³ of 2.0 Molar Sulfuric Acid	100 cm ³ of 4.0 Molar	100 cm ³
200 cm ³ of 2.4 Molar Sulfuric Acid	120 cm ³ of 4.0 Molar	80 cm ³
200 cm ³ of 2.8 Molar Sulfuric Acid	140 cm ³ of 4.0 Molar	60 cm ³
200 cm ³ of 3.2 Molar Sulfuric Acid	160 cm ³ of 4.0 Molar	40 cm ³
200 cm ³ of 3.6 Molar Sulfuric Acid	180 cm ³ of 4.0 Molar	20 cm ³
200 cm ³ of 4.0 Molar Sulfuric Acid	200 cm ³ of 4.0 Molar	0 cm ³

2.3 Risk Assessment

Name of Chemical	Source of Information	Hazards	Risks	Control Measures	Disposal Method	Emergency Procedures
Copper Sulfate (aq)	CLEAPSS Hazcards	<ul style="list-style-type: none"> • Harmful • Dangerous for the Environment 	<ul style="list-style-type: none"> • Harmful if swallowed • Irritating to eyes and skin 	<ul style="list-style-type: none"> • Do not put near mouth • Use gloves • Use goggles • Keep away from water, unless intended. 	Dissolve 64 g in 1 litre of water before pouring the solution down a foul-water drain. This disposal procedure should be kept to a minimum.	Seek medical attention. Wash contaminated area.
Hydrated Copper Sulfate (s)	CLEAPSS Hazcards	<ul style="list-style-type: none"> • Harmful • Dangerous for the Environment 	<ul style="list-style-type: none"> • Harmful if swallowed • Irritating to eyes and skin 	<ul style="list-style-type: none"> • Do not put near mouth • Use gloves • Use goggles • Label: Harmful, if above 1 mol dm⁻³ 	Crystals may be used for solutions. Dilute to less than 0.4 mol dm ⁻³ or dissolve 100 g in 1 litre of water before pouring the solution down a foul-water drain. This disposal procedure should be kept to a minimum .	Seek medical attention. Wash contaminated area.

Sulfuric Acid (aq)	CLEAPSS Hazcards	<ul style="list-style-type: none"> Corrosive, if above 1.5 mol dm⁻³ Irritant, if above 0.5 mol dm⁻³ 	Causes serious burns	<ul style="list-style-type: none"> Label: Irritant, if above 0.5 mol dm⁻³ Label: Corrosive, if above 1.5 mol dm⁻³ Wear gloves Wear goggles 	Add slowly no more than 10 cm ³ of concentrated sulfuric(VI) acid to 1 litre of 1 mol dm ⁻³ sodium carbonate solution (containing indicator) which should be constantly stirred. Let the mixture cool (or add ice), before adding more acid. Pour the solution down a foul-water drain.	Remove contaminated clothing and quickly wipe as much liquid as possible off the skin with a dry cloth before drenching the area with a large excess of water. If a large area is affected or blistering occurs, seek medical attention.
Granulated Zinc (s)	CLEAPSS Hazcards	<ul style="list-style-type: none"> Low Hazard 	N/A	Place in normal refuse	N/A	N/A

Powdered Zinc (s)	CLEAPSS Hazcards	<ul style="list-style-type: none"> • Highly Flammable • Dangerous for the environment 	When mixed with acids/alkali hydrogen is produced, flammable.	Keep away from fire.	Should be stored for removal by a Registered Waste Carrier.	Put fire out and apply cold water to any burns. Alert supervisor in the room.
Zinc Sulfate (product)	CLEAPSS Hazcards	Harmful/Irritant	Harmful if swallowed, damaged eye risk.	Wear eye protection.	Place in a suitable, closable, labelled container for disposal.	If solution is in the eye, use eye wash station for atleast 10 minutes. See a doctor.
Hydrogen (product)	CLEAPSS Hazcards	Highly Flammable.	Highly Flammable. Explosive in oxygen or chlorine.	Keep away from flames.	Burn, evaporate or vent the material in a working fume cupboard or out in the open if safe to do so.	Put fire out and apply cold water to any burns. Alert supervisor in the room.

2.3.1 Reducing Laboratory Risk

Basic Laboratory Rules

Always	Never
Wear Safety Goggles	Work in an unsupervised lab
Wear a lab coat	Eat/drink in the lab
Tie back long hair	Touch, sniff or taste chemicals
Wear gloves when handling certain chemicals	Wear open toed shoes
Use a fume cupboard if necessary	Pipette liquids by mouth
Work tidily to reduce risks	Dispose of hazardous materials down the drain
Report spillages to a supervisor	Return unused chemicals to their containers
Clear up and label containers after practical ¹⁸	

Handling Glassware

- Examine glassware for cracks/chips.
- Take care with sharp glassware.
- Correctly insert glassware into other equipment.

2.4 Methods

2.4.1 Chosen Method

Setting Up

1. Fit the bung (fitted with glass tube) into the conical flask.
2. Fit the gas syringe to the end of the glass tube.

This is shown overleaf.

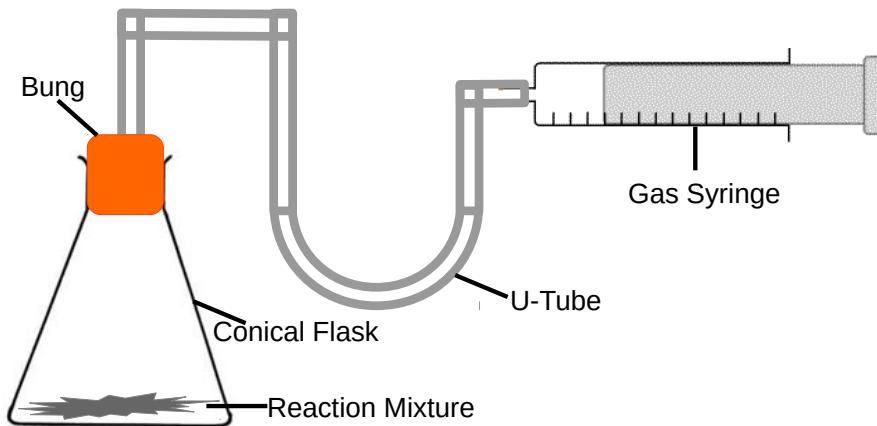


Figure 2.13: Gas Syringe Equipment

Carrying out the Experiment

1. Using a volumetric pipette measure out 40 cm^3 of the desired concentration of sulfuric acid.
2. Remove the bung from the conical flask and release the 40 cm^3 of sulfuric acid into the conical flask from the volumetric pipette.
3. Weigh out 1.0 g of granulated zinc.
4. Add the measured 1.0 g of granulated zinc to the conical flask.
5. Place the bung back in the conical flask.
6. Record the volume of hydrogen produced in cm^3 every 30 seconds for 5 minutes from the gas syringe markings to the nearest 0.5 cm^3 . Take 3 repeats of each concentration.
7. Repeat the experiment but include 10 cm^3 of copper sulfate when making up the standard solutions of sulfuric acid. Take 3 repeats of each concentration.

Interpreting the Data (as discussed before)

Volume of Hydrogen Produces vs Time Graph:

1. Plot a graph of the volume of hydrogen against time.
2. From the graph draw a tangent to the line at the initial point.
3. Calculate the gradient of the tangent by using the equation:
4. The gradient is equal to the rate of reaction.

Rate vs Concentration Graph:

1. Plot a graph of rate against concentration.
2. Determine the order of the reaction (see below).

Determining the Order of the Reaction

Once the graph of Rate vs Concentration has been drawn, the order of the reaction can be determined. Using the graph shapes shown previously you can determine the order of the reaction.

- Zero-Order Reaction - Figure 2.1 on page 10
- First-Order Reaction - Figure 2.2 on page 11
- Second-Order Reaction - Figure 2.3 on page 12

2.4.2 Justification of Chosen Method

In addition to the method discussed above, there are two other methods which would allow me to carry out my experiment:

- The Inverted Burette Method
- The Mass Change Method

Below, the methods are explained:

Inverted Burette Method

Setting up involves a similar set up to my chosen method, but the inverted burette replaces the gas syringe. This is shown overleaf:

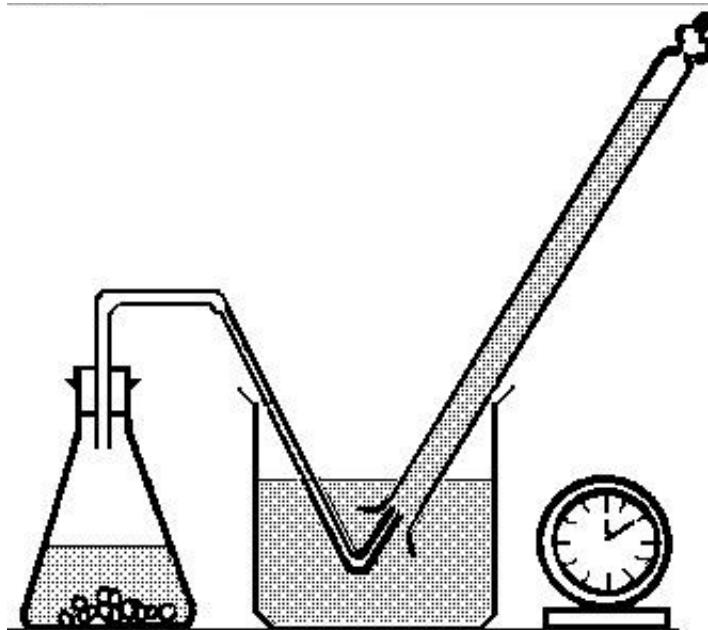


Figure 2.14: Inverted Burette Equipment

Carrying out the method is similar to my chosen method except a reading is taken from the gas syringe instead of the burette and the hydrogen is passed through a water medium. Although the gas syringes available to me have graduations every 1 ml, whereas the burettes available to me have graduations every 0.1 ml, I have still decided to carry out the gas syring method for a number of reasons. First of all the equipment is a lot easier to set up which will allow for more experiments to be carried out, thus increasing the reliability of my experiment. Secondly, as the hydrogen will not be passing through water, there is no chance of any of the hydrogen dissolving into the water, thus decreasing the validity and accuracy of my results. Finally, through my preliminary experiments, I found out that the inverted burette method produced less reliable results due to an increased range of results, as well as greater fluctuations in the results. These conclusions are backed up by my 'Preliminary Work' section of my coursework.

Mass Change Method

Setting up this method involves:

- Balance (reading to 0.01 g)
- Conical Flask

- Cotton Wool

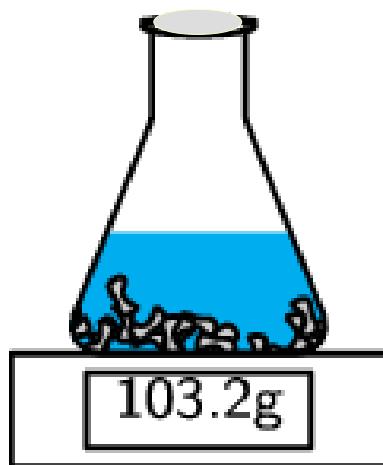


Figure 2.15: Mass Change Experiment

Carrying out this method involves:

1. Zero the Balance.
2. Pour 30 cm³ of distilled water and 10 cm³ of sulfuric acid into the conical flask.
3. Take note of the Balance value + 1.0 g
4. Weigh out 1.0 g of granulated zinc.
5. Add the measured 1.0 g of granulated zinc to the conical flask.
6. Place the cotton wool in the conical flask to stop acid 'spray' escaping.
7. Record the loss in mass in grams every 30 seconds for 5 minutes to 2 decimal places.
8. Repeat the experiment but use 30 cm³ of copper sulfate instead of distilled water.

I have chosen not to carry out this method as the balance will be a lot more sensitive to the environment and there will be room for a lot more human error. For example, left over residue could land on the scales and skew the results during the experiment. Also as hydrogen gas is produced (molecular mass of 2 per molecule) the weight of the gas which is produced would make such a little difference that it would be hard to accurately measure the mass change and actually get a significant change of mass which would produce good data to analyse and interpret.

Chapter 3

Preliminary Work

3.1 Preliminary Experiments

3.1.1 What Concentration of Sulfuric Acid to use?

Originally I planned to use 10 cm³ of sulfuric acid and then add 30 cm³ of distilled water, but I found that this was too slow. Therefore I decided to use 40 cm³ of sulfuric acid and not dilute with any further water. From my experiments I decided that 2 molar sulfuric acid would be the starting concentration for my range of experiments without a catalyst.

3.1.2 Which Form of Zinc to use?

I ordered two types of zinc in order to compare the use of them. These two forms were granulated and powdered zinc. Below is a picture of the granulated zinc.



Figure 3.1: Granulated Zinc.

As you can see the sizes of the chunks of the zinc vary dramatically. This meant that I could not accurately measure 1.0 g of zinc using the granulated zinc, and also the varying surface area to volume ratio would cause the experiment to not be accurate and produce unreliable results. The powdered zinc was extremely finely grinded which allowed for a more reliable surface area to volume ratio, which as discussed in the chemical ideas section has an impact on the rate of reaction.

3.1.3 Inverted Burette Method vs Gas Syringe Method

To work out which method would be best to carry out my experiment, I decided to carry out the two most viable methods. From this experiment I decided that the method which gave the most reliable results would be the method I used to carry out my actual experiment. For this experiment I decided to measure the volume of hydrogen gas produced over a 5 minute period, with 1.0 g of powdered zinc and 40 cm³ of 2.0 molar sulfuric acid. My results are shown below for both method.

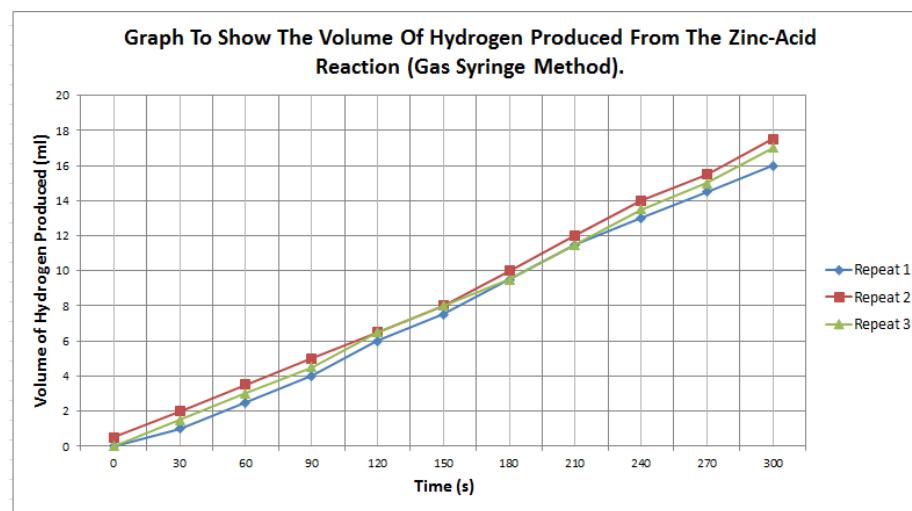


Figure 3.2: Gas Syringe Method Graph

Above shows the graph of my results for the gas syringe method. This shows that the average volume collected was 16.7 ml of gas and the average range was only 1.0 ml. For the raw data table which I used to analyse this from please look at Figure 3.8 on page 45 in the preliminary work section appendix.

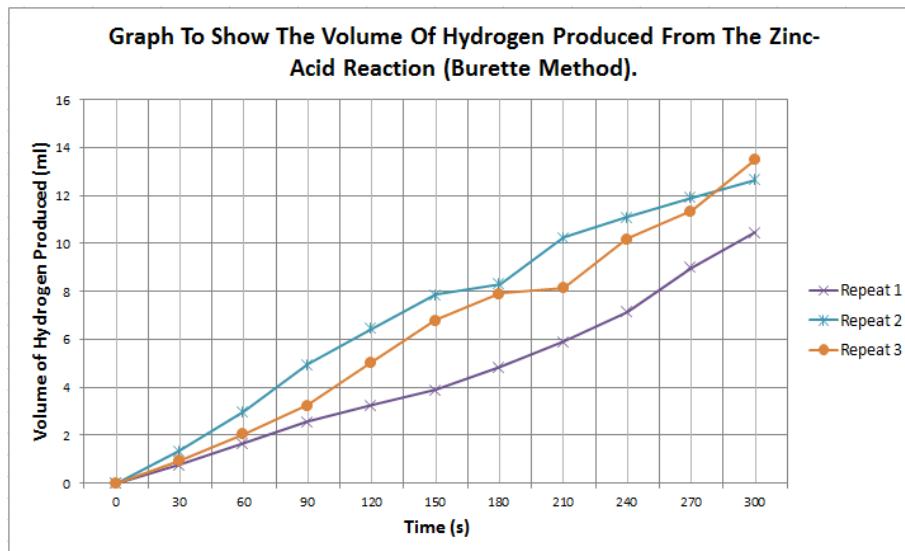


Figure 3.3: Burette Method Graph

Above shows the graph of my results for the burette method. This shows that the average volume collected was 11.87 ml and the average range was 2.20 ml. For the raw data table which I used to analyse this from please look at Figure 3.9 on page 46 in the preliminary work section appendix.

In conclusion I decided that the gas syringe method was more reliable as the average range was only 1.0 ml in comparison to the inverted burette method which produced a significantly larger range of 2.2 ml. The burette method also collected significantly less hydrogen. I believe this occurred as the hydrogen was being passed through a water medium, therefore some of the gas was absorbed into the water. This also was a factor that made me chose the gas syringe method over the burette method, as in the gas syringe method the hydrogen produced does not pass through a water medium and therefore would not dissolve into it. Although the burette read to two decimal places instead of one, the actual results produced by the burette method appeared to be unreliable, therefore I have chosen to use the gas syringe method for my experiment.

3.1.4 Measuring the Temperature of the Reaction

As room temperature varies from day to day I decided to measure the the temperature at the beginning of the day of experiments and as the Zinc-Acid reaction is an exothermic reaction, I knew that the temperature of the solution would increase as time went on. It turned out that all of the days that I carried

out my experiments the solutions were all at the same temperature. They also produced the same temperature increase of 1°C. This graph is shown below.

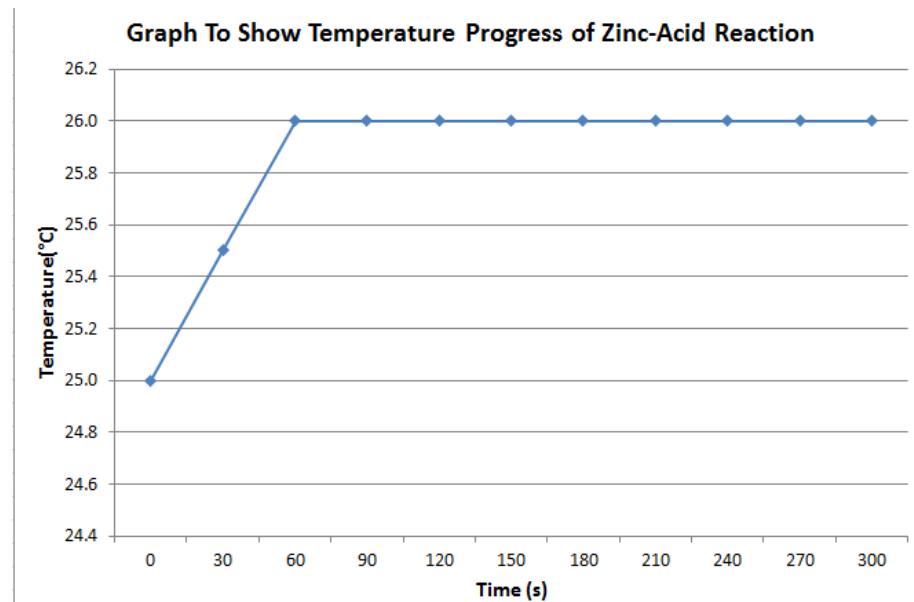


Figure 3.4: Graph To Show The Temperature Progress of the Zinc-Acid Reaction.

Below you can see the apparatus that I used in order to measure the temperature of the reaction.

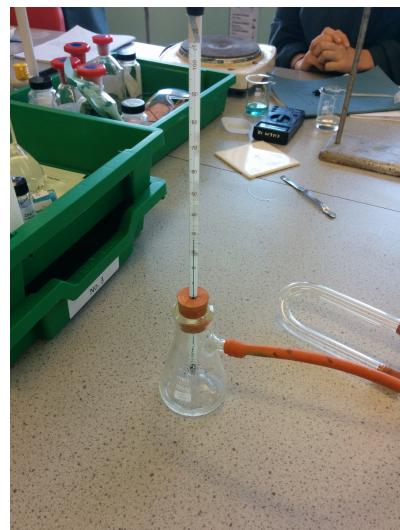


Figure 3.5: Thermometer set-up.

My set-up involved the bung in the conical flask being modified to have a hole in it to stick a thermometer in it. Although the set-up allowed me to measure the temperature of the reaction, I decided not to measure the temperature of every experiment series as the hole in the bung caused the set-up to not be air tight. This meant that I would not collect the correct amount of hydrogen produced from the reaction. Below the two images show the set-up being in place, but when the gas syringe is pushed in, an excess amount of volume is lost.

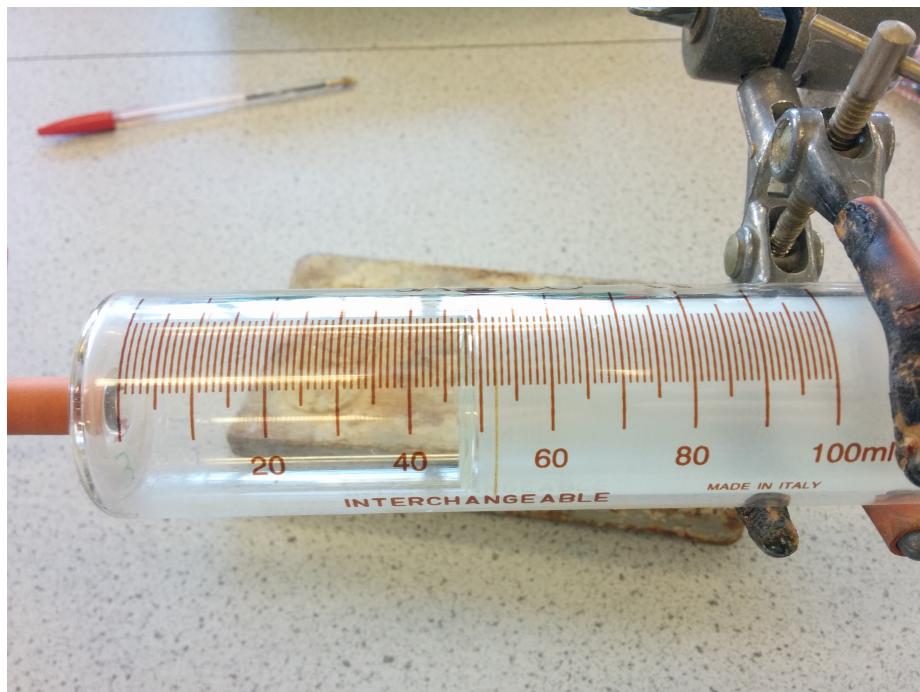


Figure 3.6: Before Pushing The Gas Syringe.

The image above shows the value of 52.0 ml of gas before pushing the gas syringe in.

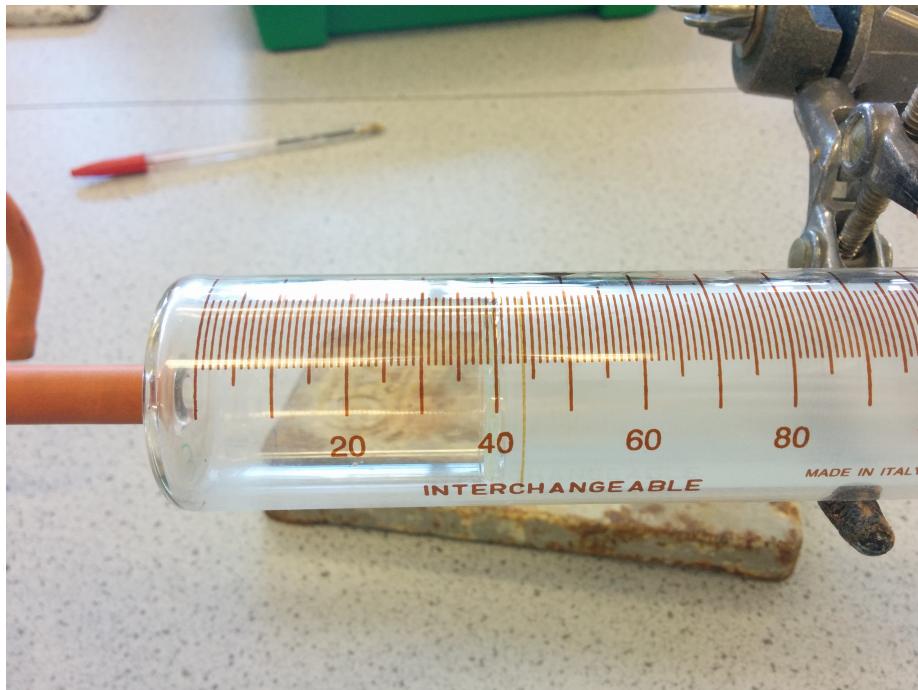


Figure 3.7: After Pushing The Gas Syringe.

The image below shows the value of 44.0 ml of gas after pushing the gas syringe in. This means that a total of 8.0 ml of gas was lost from pushing the gas syringe in, this indicates a significant gas leak. Therefore I used a complete bung and this shows to be air tight, as when it was pushed in the gas syringe returned to the same value. For the raw data table which I used to analyse this from please look at Figure 3.10 on page 46 in the preliminary work section appendix.

3.2 Techniques

Through carrying out my preliminary work, I found out a number of techniques that I could use to increase the accuracy and validity of my results.

3.2.1 Order of Adding Reactants

Through my preliminary experiments I found that adding the sulfuric acid first then the zinc caused the zinc to have a tendency to form clumps, thus decreasing the volume of hydrogen produced due to a decreased surface area. Therefore in

my real experiment I will add the zinc first then pour the sulfuric acid into the conical flask after.

3.2.2 Spreading the Zinc

As zinc powder has a tendency to form clumps when a solution is added, I decided to spread the zinc across the bottom of the conical flask before adding the sulfuric acid. This reduced the clumping effect of the zinc powder and as I spread the zinc in the same way every time, it led to a more reliable and accurate experiment.

3.2.3 Cleaning the Gas Syringe

I found that sometimes (due to others using the same equipment in between me) the gas syringe had something in it which caused it to get stuck at certain points. Because of this I decided to clean out the gas syringe every time I used it to reduce the friction and chance of it getting stuck.

3.3 Preliminary Work Raw Data Appendix

3.3.1 Gas Syringe Method Raw Data

Time (s)	Volume Of Hydrogen Produced (ml)			
	Repeat 1	Repeat 2	Repeat 3	Average
0	0.0	0.5	0.0	0.2
30	1.0	2.0	1.5	1.5
60	2.5	3.5	3.0	3.0
90	4.0	5.0	4.5	4.5
120	6.0	6.5	6.5	6.3
150	7.5	8.0	8.0	7.8
180	9.5	10.0	9.5	9.7
210	11.5	12.0	11.5	11.7
240	13.0	14.0	13.5	13.5
270	14.5	15.5	15.0	15.0
300	16.0	17.5	17.0	16.8

Figure 3.8: Gas Syringe Method Raw Data

3.3.2 Burette Method Raw Data

Time (s)	Volume of Hydrogen Produced (ml)			
	Repeat 1	Repeat 2	Repeat 3	Average
0	0.00	0.00	0.00	0.00
30	0.75	1.35	0.95	1.02
60	1.65	3.00	2.05	2.23
90	2.55	4.95	3.25	3.58
120	3.25	6.45	5.05	4.92
150	3.90	7.85	6.80	6.18
180	4.85	8.30	7.90	7.02
210	5.90	10.25	8.15	8.10
240	7.15	11.10	10.20	9.48
270	9.00	11.90	11.35	10.75
300	10.45	12.65	13.50	12.20

Figure 3.9: Burette Method Raw Data

3.3.3 Temperature Experiment Raw Data

Time (s)	Temp (°C)
0	25.0
30	25.5
60	26.0
90	26.0
120	26.0
150	26.0
180	26.0
210	26.0
240	26.0
270	26.0
300	26.0

Figure 3.10: Temperature Raw Data

Chapter 4

Experiment

All of my experiments are carried out using my chosen method discussed in the 'Methods' section on page 33. The columns highlighted in red are outliers that I will not be using in my 'Analysis and Interpretation' section. All of my experiments are also carried out using 40 cm³ of solution with 1.0 g of zinc, unless otherwise stated.

4.1 Non-Catalysed Experiment Series

In order to work out the order of sulfuric acid in the Zinc and sulfuric acid reaction I carried out an experiment without the presence catalyst. I varied the concentration of the sulfuric acid (6 series with a range of 2.0 Molar - 6.0 Molar) whilst keeping the mass of zinc the same (1.0 g). To make up my solutions I followed the procedure I outlined in my 'Preparing Chemicals' section on page 27. My raw data tables can be seen below for each concentration of sulfuric acid.

4.1.1 2.0 Molar Sulfuric Acid

Time (s)	Volume of Hydrogen Produced (ml)			Average
	Repeat 1	Repeat 2	Repeat 3	
0	0.0	0.5	0.0	0.2
30	1.0	2.0	1.5	1.5
60	2.5	3.5	3.0	3.0
90	4.0	5.0	4.5	4.5
120	6.0	6.5	6.5	6.3
150	7.5	8.0	8.0	7.8
180	9.5	10.0	9.5	9.7
210	11.5	12.0	11.5	11.7
240	13.0	14.0	13.5	13.5
270	14.5	15.5	15.0	15.0
300	16.0	17.5	17.0	16.8

Figure 4.1: 2 Molar Sulfuric Acid and 1.0 g of Zinc

4.1.2 2.4 Molar Sulfuric Acid

Time (s)	Volume of Hydrogen Produced (ml)			Average
	Repeat 1	Repeat 2	Repeat 3	
0	0.5	0.5	0.0	0.3
30	3.0	3.5	3.0	3.2
60	5.0	6.5	5.5	5.7
90	7.5	9.0	8.0	8.2
120	9.5	12.0	10.5	10.7
150	13.0	15.5	13.5	14.0
180	16.0	19.0	16.5	17.2
210	19.0	22.0	19.5	20.2
240	22.0	25.0	22.5	23.2
270	25.5	28.0	25.5	26.3
300	28.5	30.5	28.5	29.2

Figure 4.2: 2.4 Molar Sulfuric Acid and 1.0 g of Zinc

4.1.3 2.8 Molar Sulfuric Acid

Time (s)	Volume of Hydrogen Produced (ml)			
	Repeat 1	Repeat 2	Repeat 3	Average
0	1.0	0.5	0.5	0.7
30	5.0	4.0	6.0	5.0
60	10.0	6.0	13.5	9.8
90	18.0	9.5	24.5	17.3
120	24.0	13.0	31.0	22.7
150	30.0	18.5	37.0	28.5
180	36.0	26.0	43.0	35.0
210	41.0	33.5	48.0	40.8
240	47.0	39.5	52.0	46.2
270	53.0	43.5	56.0	50.8
300	57.5	48.0	59.5	55.0

Figure 4.3: 2.8 Molar Sulfuric Acid and 1.0 g of Zinc

4.1.4 3.2 Molar Sulfuric Acid

Time (s)	Volume of Hydrogen Produced (ml)			
	Repeat 1	Repeat 2	Repeat 3	Average
0	1.5	1.0	1.5	1.3
30	10.0	14.0	12.0	12.0
60	25.0	27.0	26.0	26.0
90	38.0	39.0	38.5	38.5
120	47.0	47.0	48.5	47.5
150	56.0	55.0	56.5	55.8
180	64.0	63.5	64.0	63.8
210	70.5	70.0	71.0	70.5
240	80.0	78.5	79.5	79.3
270	84.5	84.0	84.5	84.3
300	91.5	90.0	91.0	90.8

Figure 4.4: 3.2 Molar Sulfuric Acid and 1.0 g of Zinc

4.1.5 3.6 Molar Sulfuric Acid

Time (s)	Volume of Hydrogen Produced (ml)			Average
	Repeat 1	Repeat 2	Repeat 3	
0	1.0	1.0	0.5	0.8
30	6.0	8.5	9.0	7.8
60	15.0	18.0	17.5	16.8
90	26.0	29.5	29.5	28.3
120	34.5	38.0	38.5	37.0
150	42.0	46.5	47.0	45.2
180	49.5	55.0	56.0	53.5
210	59.5	64.0	63.5	62.3
240	67.5	73.5	73.5	71.5
270	77.5	81.0	81.5	80.0
300	87.0	89.5	89.0	88.5

Figure 4.5: 3.6 Molar Sulfuric Acid and 1.0 g of Zinc

4.1.6 4.0 Molar Sulfuric Acid

Time (s)	Volume of Hydrogen Produced (ml)			Average
	Repeat 1	Repeat 2	Repeat 3	
0	1.0	1.0	1.0	1.0
30	12.0	6.5	10.5	9.7
60	27.0	15.0	22.0	21.3
90	37.0	24.5	31.5	31.0
120	42.0	36.0	42.0	40.0
150	53.5	47.0	52.0	50.8
180	60.5	59.5	59.5	59.8
210	67.0	67.0	68.5	67.5
240	73.5	74.0	79.0	75.5
270	78.0	77.5	84.5	80.0
300	84.5	86.0	90.0	86.8

Figure 4.6: 4.0 Molar Sulfuric Acid and 1.0 g of Zinc

4.2 Copper Sulfate Catalysed Experiment Series

My previous experiment series was carried out so that I would be able to work out the order of sulfuric acid without the presence of a catalyst; however the acid may be a different order in presence of a catalyst. For this experiment series I have created 40 cm³, 0.01 molar copper sulfate and a varying concentration of sulfuric acid (6 series from 0.2 molar - 1.2 molar). I prepared these solutions by adding the desired amount of sulfuric acid to make the solution and then adding 10 cm³ of 0.1 molar copper sulfate (which I prepared, as outlined in the 'Preparing Chemicals' section on page 27), both using a volumetric pipette to carry out the measurements. After the two chemicals were in the volumetric flask I filled the volumetric flask with distilled water up until the meniscus of the solution was on the line. My raw data tables are shown below for this experiment.

4.2.1 0.2 Molar Sulfuric Acid

Time (s)	Volume of Hydrogen Produced (ml)			Average
	Repeat 1	Repeat 2	Repeat 3	
0	1.0	1.0	2.0	1.0
10	8.5	8.0	10.0	8.3
20	14.0	15.0	15.5	14.5
30	17.0	18.0	20.0	17.5
40	20.0	20.5	25.5	20.3
50	22.5	22.5	30.0	22.5
60	24.0	24.0	32.5	24.0
70	25.0	25.5	36.0	25.3
80	26.0	26.5	38.0	26.3
90	27.5	27.5	40.0	27.5
100	29.0	28.5	42.5	28.8

Figure 4.7: 0.2 Molar Sulfuric Acid, 0.01 Molar Copper Sulfate and 1.0 g of Zinc

4.2.2 0.4 Molar Sulfuric Acid

Time (s)	Volume of Hydrogen Produced (ml)			Average
	Repeat 1	Repeat 2	Repeat 3	
0	4.0	3.5	4.0	3.8
10	19.0	20.5	20.0	19.8
20	27.0	28.0	27.5	27.5
30	33.0	34.0	33.5	33.5
40	39.0	39.5	39.5	39.3
50	44.0	45.0	45.0	44.7
60	49.5	50.0	50.0	49.8
70	52.5	55.0	53.5	53.7
80	56.5	59.0	57.0	57.5
90	60.5	62.0	61.0	61.2
100	63.5	65.5	64.5	64.5

Figure 4.8: 0.4 Molar Sulfuric Acid, 0.01 Molar Copper Sulfate and 1.0 g of Zinc

4.2.3 0.6 Molar Sulfuric Acid

Time (s)	Volume of Hydrogen Produced (ml)			Average
	Repeat 1	Repeat 2	Repeat 3	
0	3.5	7.0	5.0	6.0
10	9.0	20.0	18.0	19.0
20	18.5	27.0	25.5	26.3
30	26.0	33.5	30.5	32.0
40	35.0	38.5	33.5	36.0
50	40.0	42.0	41.5	41.8
60	43.5	45.5	46.0	45.8
70	47.0	48.0	48.5	48.3
80	49.5	52.0	51.0	51.5
90	52.5	54.5	53.0	53.8
100	55.5	59.0	56.0	57.5

Figure 4.9: 0.6 Molar Sulfuric Acid, 0.01 Molar Copper Sulfate and 1.0 g of Zinc

4.2.4 0.8 Molar Sulfuric Acid

Time (s)	Volume of Hydrogen Produced (ml)			
	Repeat 1	Repeat 2	Repeat 3	Average
0	6.0	6.0	6.0	6.0
10	30.0	27.5	28.0	28.5
20	45.0	39.5	41.0	41.8
30	54.0	46.0	48.5	49.5
40	60.5	52.0	55.0	55.8
50	66.0	57.5	60.0	61.2
60	70.0	62.0	65.5	65.8
70	75.0	67.5	70.5	71.0
80	86.0	73.0	76.0	78.3
90	84.5	77.0	80.0	80.5
100	88.0	80.0	85.0	84.3

Figure 4.10: 0.8 Molar Sulfuric Acid, 0.01 Molar copper sulfate and 1.0 g of Zinc

4.2.5 1.0 Molar Sulfuric Acid

Time (s)	Volume of Hydrogen Produced (ml)			
	Repeat 1	Repeat 2	Repeat 3	Average
0	8.0	8.0	8.0	8.0
10	29.0	30.0	29.5	29.5
20	40.0	41.0	40.0	40.3
30	51.0	50.5	51.0	50.8
40	60.0	61.0	60.0	60.3
50	65.5	66.0	66.0	65.8
60	70.5	72.0	70.5	71.0
70	77.0	80.0	77.5	78.2
80	80.5	85.0	82.5	82.7
90	85.5	90.5	87.5	87.8
100	88.5	93.5	90.0	90.7

Figure 4.11: 1.0 Molar Sulfuric Acid, 0.01 Molar Copper Sulfate and 1.0 g of Zinc

4.2.6 1.2 Molar Sulfuric Acid

Time (s)	Volume of Hydrogen Produced (ml)			Average
	Repeat 1	Repeat 2	Repeat 3	
0	9.0	9.0	9.0	9.0
10	27.0	35.0	30.0	32.5
20	33.0	50.0	40.0	45.0
30	37.0	58.0	52.0	55.0
40	41.0	68.0	60.5	64.3
50	46.0	70.5	66.5	68.5
60	50.0	77.0	73.0	75.0
70	55.5	82.0	77.5	79.8
80	59.5	87.0	82.0	84.5
90	65.5	90.0	85.5	87.8
100	68.0	95.5	90.5	93.0

Figure 4.12: 1.2 Molar Sulfuric Acid, 0.01 Molar Copper Sulfate and 1.0 g of Zinc

4.3 Varying Copper Sulfate Experiment Series

As I need to find the order of copper sulfate in the reaction, the next logical experiment was to change the concentration of copper sulfate whilst keeping the sulfuric acid concentration the same. For this experiment I used 0.2 molar sulfuric acid with a range of copper sulfate concentrations (6 series from 0.01 molar - 0.06 molar). I made these solutions up the same way as before, but instead of varying the volume of sulfuric acid I initially added to the volumetric flask, I varied the volume of copper sulfate solution I added. My raw data tables are shown below for this experiment.

4.3.1 0.01 Molar Copper Sulfate

Please see Figure 4.7 on page 51 for the raw data table of this experiment series.

4.3.2 0.02 Molar Copper Sulfate

Time (s)	Volume of Hydrogen Produced (ml)			
	Repeat 1	Repeat 2	Repeat 3	Average
0	2.0	1.5	1.5	1.7
10	9.0	8.0	8.5	8.5
20	14.0	14.5	15.0	14.5
30	18.0	19.0	19.5	18.8
40	21.0	22.0	22.5	21.8
50	24.0	24.5	25.0	24.5
60	27.0	27.5	28.0	27.5
70	29.5	30.0	31.0	30.2
80	32.0	33.0	33.5	32.8
90	34.0	35.0	36.0	35.0
100	36.0	37.0	37.5	36.8

Figure 4.13: 0.2 Molar Sulfuric Acid, 0.02 Molar Copper Sulfate and 1.0 g of Zinc

4.3.3 0.03 Molar Copper Sulfate

Time (s)	Volume of Hydrogen Produced (ml)			
	Repeat 1	Repeat 2	Repeat 3	Average
0	3.0	3.0	3.0	3.0
10	11.5	12.0	11.5	11.7
20	16.0	16.5	17.0	16.5
30	21.5	22.0	22.0	21.8
40	25.0	27.0	25.5	25.8
50	28.0	29.0	28.0	28.3
60	30.0	31.0	31.0	30.7
70	34.5	35.0	34.0	34.5
80	37.0	38.0	37.0	37.3
90	39.0	40.0	40.0	39.7
100	42.0	42.5	43.0	42.5

Figure 4.14: 0.2 Molar Sulfuric Acid, 0.03 Molar Copper Sulfate and 1.0 g of Zinc

4.3.4 0.04 Molar Copper Sulfate

Time (s)	Volume of Hydrogen Produced (ml)			
	Repeat 1	Repeat 2	Repeat 3	Average
0	3.5	3.5	4.0	3.7
10	13.5	13.0	13.0	13.2
20	18.5	17.5	18.0	18.0
30	25.0	23.5	24.0	24.2
40	29.0	28.0	28.0	28.3
50	33.0	34.0	33.0	33.3
60	36.0	37.0	37.0	36.7
70	39.5	40.0	40.5	40.0
80	42.0	43.0	44.0	43.0
90	45.5	46.0	47.0	46.2
100	48.0	49.0	50.0	49.0

Figure 4.15: 0.2 Molar Sulfuric Acid, 0.04 Molar Copper Sulfate and 1.0 g of Zinc

4.3.5 0.05 Molar Copper Sulfate

Time (s)	Volume of Hydrogen Produced (ml)			
	Repeat 1	Repeat 2	Repeat 3	Average
0	4.0	4.5	4.0	4.2
10	17.0	17.0	18.0	17.3
20	23.5	24.0	23.5	23.7
30	27.5	28.0	28.5	28.0
40	30.5	31.0	32.0	31.2
50	34.5	35.0	36.0	35.2
60	38.0	39.0	42.0	39.7
70	42.0	42.5	48.0	44.2
80	45.0	48.5	51.0	48.2
90	58.0	52.0	53.0	54.3
100	53.5	55.0	58.0	55.5

Figure 4.16: 0.2 Molar Sulfuric Acid, 0.05 Molar Copper Sulfate and 1.0 g of Zinc

4.3.6 0.06 Molar Copper Sulfate

Time (s)	Volume of Hydrogen Produced (ml)			Average
	Repeat 1	Repeat 2	Repeat 3	
0	4.5	4.5	5.0	4.7
10	16.5	16.0	17.0	16.5
20	25.0	25.0	26.0	25.3
30	32.0	32.0	31.0	31.7
40	36.0	38.0	37.5	37.2
50	41.0	41.5	42.5	41.7
60	45.5	46.0	47.0	46.2
70	50.0	51.0	52.0	51.0
80	54.5	55.0	56.0	55.2
90	59.0	59.0	60.0	59.3
100	63.5	64.0	65.0	64.2

Figure 4.17: 0.2 Molar Sulfuric Acid, 0.06 Molar Copper Sulfate and 1.0 g of Zinc

4.4 Different Catalysts Experiment Series

After finishing my initial experiment, I began to think about how different catalysts affect the rate of reaction. As my college had a low stock of chemicals by the time I began this experiment, I could only get 4 different metal sulfates. These are listed below.

- Copper Sulfate
- Nickel Sulfate
- Manganese Sulfate
- Iron Sulfate

I made up the solutions using 0.4 molar sulfuric acid and 0.01 molar of the catalyst. These were made up using the same method as I used for the rest of my solutions. This can be found in the 'Preparing Chemicals' section on page 27.

4.4.1 Copper Sulfate

As the rest of my experiments in this series only have one repeat, I have decided to take the average of the 3 repeats from Figure 4.8 on page 52 for the table in

this section.

Time(s)	Volume of Hydrogen Produced (ml)
0	0.0
10	16.0
20	23.7
30	29.7
40	35.5
50	40.8
60	46.0
70	49.8
80	53.7
90	57.3
100	60.7

Figure 4.18: 0.4 Molar Sulfuric Acid, 0.01 Molar Copper Sulfate and 1.0 g of Zinc

4.4.2 Nickel Sulfate

Time(s)	Volume of Hydrogen Produced (ml)
0	2.0
10	10.0
20	15.5
30	20.5
40	24.5
50	28.0
60	32.0
70	35.0
80	38.0
90	41.5
100	44.0

Figure 4.19: 0.4 Molar Sulfuric Acid, 0.01 Molar Nickel Sulfate and 1.0 g of Zinc

4.4.3 Manganese Sulfate

Time(s)	Volume of Hydrogen Produced (ml)
0	3.0
10	11.5
20	18.0
30	25.0
40	30.0
50	36.0
60	42.0
70	46.5
80	51.0
90	56.0
100	60.0

Figure 4.20: 0.4 Molar Sulfuric Acid, 0.01 Molar Manganese Sulfate and 1.0 g of Zinc

4.4.4 Iron Sulfate

Time(s)	Volume of Hydrogen Produced (ml)
0	2.5
10	12.0
20	18.0
30	23.0
40	26.5
50	30.0
60	33.5
70	36.0
80	39.0
90	43.0
100	46.5

Figure 4.21: 0.4 Molar Sulfuric Acid, 0.01 Molar Iron Sulfate and 1.0 g of Zinc

4.5 Solubility of Hydrogen Experiment Series

As shown by my non-catalyst experiment, it appears that the 3.2 molar experiment (Figure 4.4 on page 49) has a faster rate, and produces more hydrogen gas over 5 minutes, I began to wonder about the solubility of hydrogen within different acid concentrations. I predicted that the higher the concentration of acid, the higher the solubility of hydrogen. For this experiment I made sulfuric acid the excess reactant by using 0.1 g of zinc and 40 cm³ of a range of different concentrations of sulfuric acid. The solutions used in this experiment were made up using the same method as I used for the rest of my solutions. This can be found in the 'Preparing Chemicals' section on page 27.

4.5.1 1 Molar Sulfuric Acid

Time(s)	Volume of Hydrogen Produced (ml)
0	0.0
30	1.0
60	2.0
90	3.0
120	3.5
150	4.0
180	4.5
210	5.0
240	5.5
270	6.0
300	6.5

Figure 4.22: 1.0 Molar Sulfuric Acid and 0.1 g of Zinc

4.5.2 2 Molar Sulfuric Acid

Time(s)	Volume of Hydrogen Produced (ml)
0	0.0
30	1.5
60	2.5
90	3.5
120	4.0
150	4.5
180	5.0
210	5.5
240	6.0
270	6.5
300	7.0

Figure 4.23: 2.0 Molar Sulfuric Acid and 0.1 g of Zinc

4.5.3 3 Molar Sulfuric Acid

Time(s)	Volume of Hydrogen Produced (ml)
0	0.0
30	2.0
60	3.0
90	4.0
120	4.5
150	5.0
180	5.5
210	6.0
240	6.5
270	6.5
300	7.0

Figure 4.24: 3.0 Molar Sulfuric Acid and 0.1 g of Zinc

4.5.4 4 Molar Sulfuric Acid

Time(s)	Volume of Hydrogen Produced (ml)
0	0.0
30	2.0
60	3.0
90	4.0
120	4.5
150	5.0
180	5.5
210	6.0
240	6.5
270	6.5
300	6.5

Figure 4.25: 4.0 Molar Sulfuric Acid and 0.1 g of Zinc

Chapter 5

Analysis and Interpretation

All of the graphs are interpreted from raw data tables shown in the Experiment section of the coursework. Straight line graphs gradients are worked out from the straight line, and curved lines gradients are worked out by a red line extending from $t = 0$. The calculations are then shown below each graph. The graphs plotted take into account the fact that when the bung is pushed into the conical flask the volume of the system decreases which therefore increases the pressure and falsely reads on the gas syringe that hydrogen gas has been produced. Therefore the $Time = 0$ value for each data series has been taken away from every value, thus causing each series to pass through the origin.

5.1 Zinc and Sulfuric Acid Without a Catalyst

Below are the graphs that I have used to work out the rate of the reaction for the experiment series that did not involve a catalyst. All graphs have taken a mean average of the three repeats in order to plot the data points. As discussed in my 'Chosen Method' section on page 33 I am using the equation $\text{Gradient} = \frac{\text{Change in } Y}{\text{Change in } X}$ and as the gradient is equal to the rate of reaction, I will be able to use the graphs below in order to generate a rate vs concentration graph which will determine the order of the reaction. I will calculate all rates to 3 decimal places.

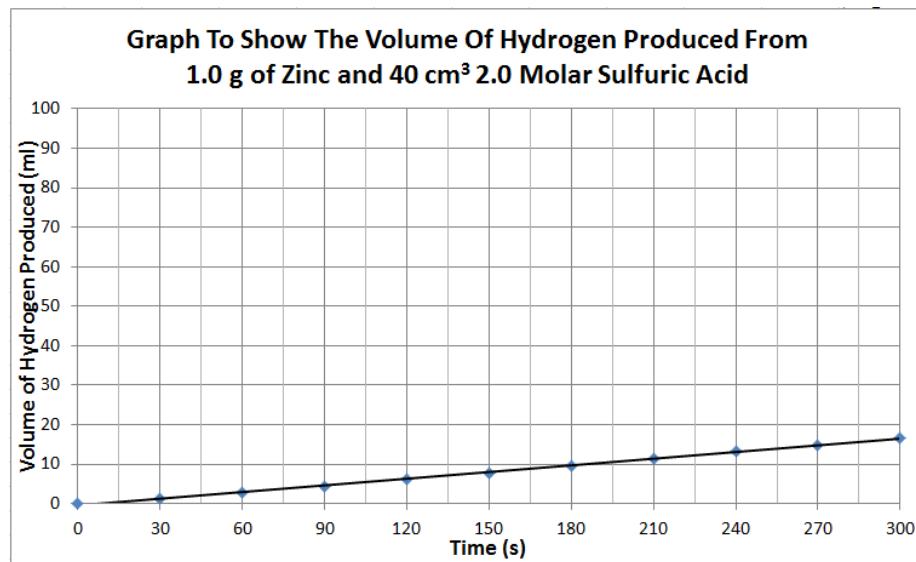


Figure 5.1: 2.0 Molar Sulfuric Acid and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{16.67}{300}$$

Therefore:

$$\text{Rate} = 0.056 \text{ ml s}^{-1}$$

This graph's shape is similar to the majority of the graphs that I have produced. This steady, straight line can be explained by the fact that there is still the maximum amount of reactants available to react over the period of time that my experiment was taken. Eventually, if the graph continued, you would see the graph start to curve off due to less reactant molecules being available to react and therefore less likely for molecules to collide per unit time.

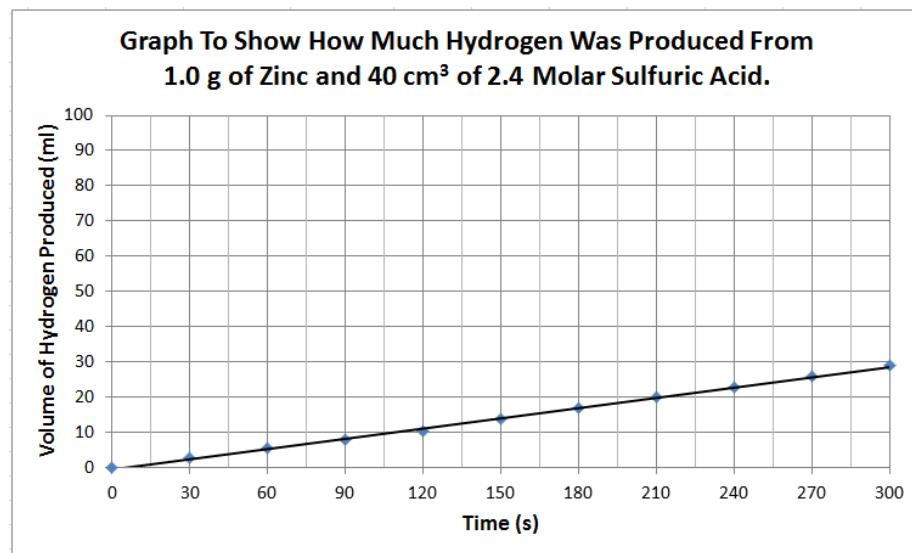


Figure 5.2: 2.4 Molar Sulfuric Acid and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{28.83}{300}$$

Therefore:

$$\text{Rate} = 0.096 \text{ ml s}^{-1}$$

**Graph To Show How Much Hydrogen Was Produced From 1.0 g
of Zinc and 40 cm³ of 2.8 Molar Sulfuric Acid.**

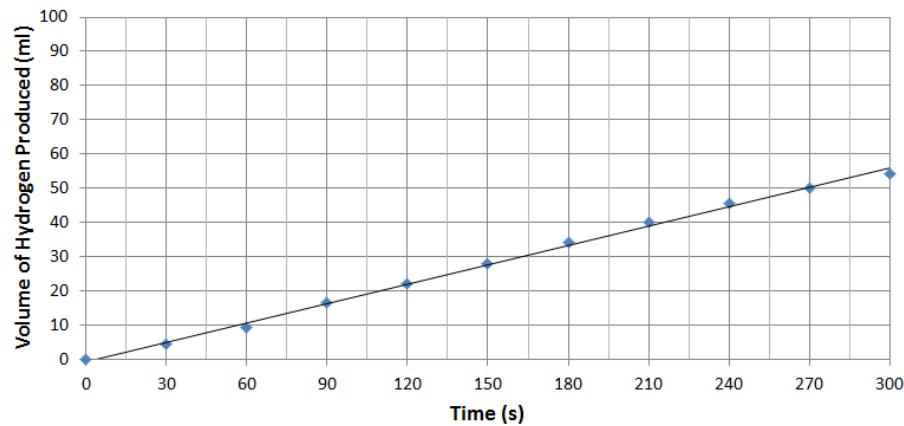


Figure 5.3: 2.8 Molar Sulfuric Acid and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{55.50}{300}$$

Therefore:

$$\text{Rate} = 0.185 \text{ ml s}^{-1}$$

Graph To Show How Much Hydrogen Was Produced From 1.0 g of Zinc and 40 cm³ of 3.2 Molar Sulfuric Acid.

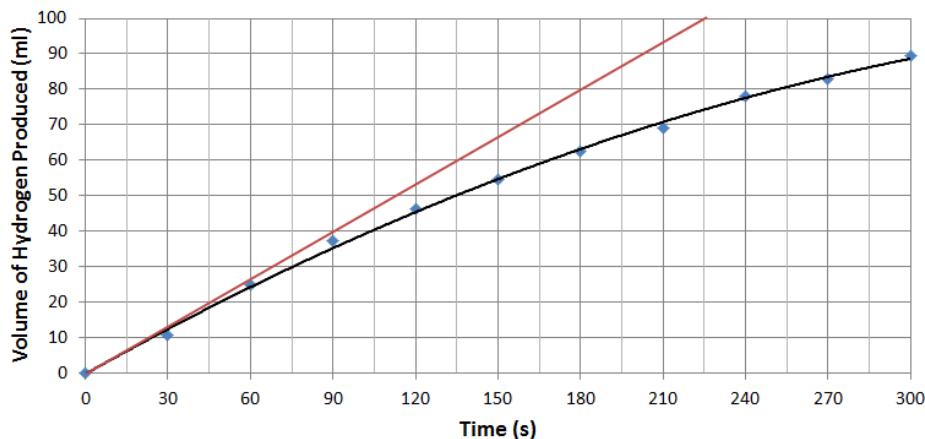


Figure 5.4: 3.2 Molar Sulfuric Acid and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{100}{227.14}$$

Therefore:

$$\text{Rate} = 0.440 \text{ ml s}^{-1}$$

This graph appears to start curving off as time goes on, thus meaning that the reaction gets slower over time. This shape can be explained by the fact that as the reaction progresses, there are less reactants available to react and therefore as stated by collision theory, the rate of reaction will decrease. This is discussed in more detail in section 2.1.2 on page 5. For a reaction to occur, molecules must collide with enough kinetic energy in order to overcome the repulsive and bonding forces of the molecule. Therefore as there will be less reactants in a given amount of space, there will be less chance of molecules colliding, thus making the reaction less likely as time goes on.

Graph To Show How Much Hydrogen Was Produced From 1.0 g of Zinc and 40 cm³ of 3.6 Molar Sulfuric Acid.

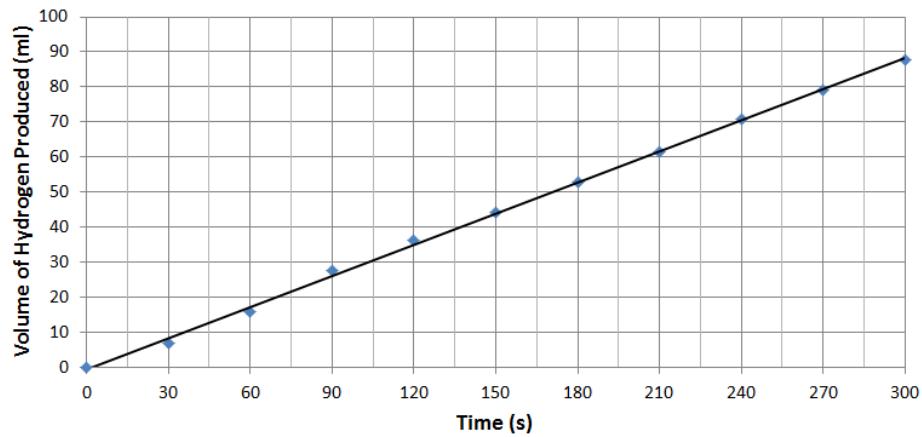


Figure 5.5: 3.6 Molar Sulfuric Acid and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{87.67}{300}$$

Therefore:

$$\text{Rate} = 0.292 \text{ ml s}^{-1}$$

Graph To Show How Much Hydrogen Was Produced From 1.0 g of Zinc and 40 cm³ of 4.0 Molar Sulfuric Acid.

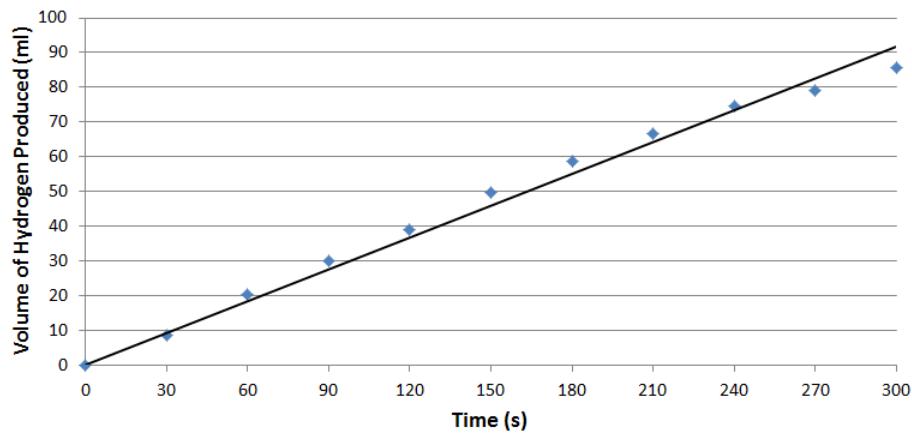


Figure 5.6: 4.0 Molar Sulfuric Acid and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{92}{300}$$

Therefore:

$$\text{Rate} = 0.307 \text{ ml s}^{-1}$$

Below the rates of reactions are summarised in a table.

Sulfuric Acid Concentration (Molar)	Initial Rate of Reaction (ml s ⁻¹)
2.0	0.056
2.4	0.096
2.8	0.185
3.2	0.440
3.6	0.292
4.0	0.307

Figure 5.7: Initial Rates of Reactions for the Non-Catalysed Zinc and Sulfuric Acid Reaction

Below the rates of reactions are compiled into a rate vs concentration graph.

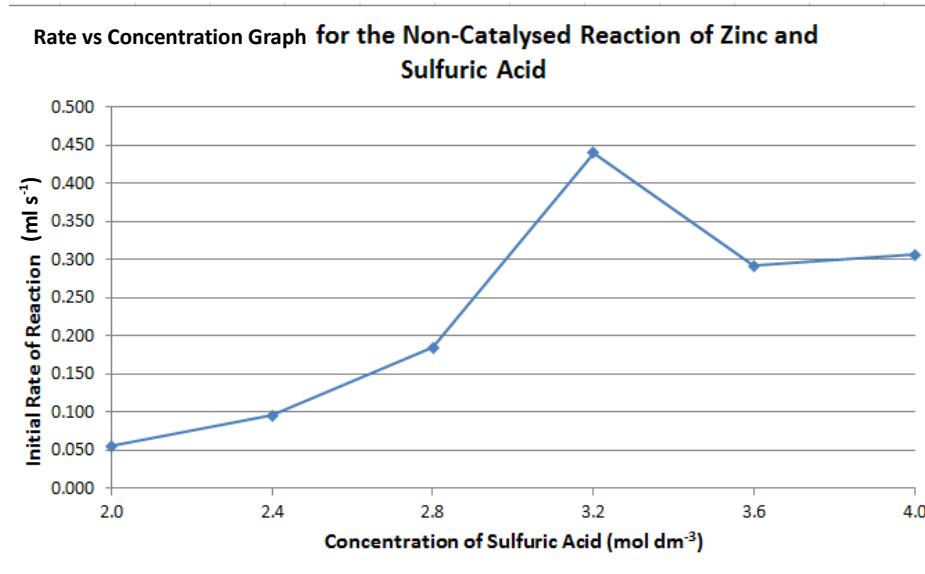


Figure 5.8: Rate vs Concentration Graph for the Non-Catalysed Reaction

The rate vs concentration graph for this experiment appears to be a second order reaction up to 3.2 Molar; however I believe that the zinc clumps at higher concentrations of acids, which would explain the decreased rate of reaction after 3.2 molar. Therefore I think that the order of sulfuric acid without the presence of a catalyst is second order. This means that doubling the concentration of sulfuric acid in this reaction quadruples the initial rate of reaction. The rate equation for this reaction would be:

$$\text{Rate} = k[\text{Sulfuric Acid}]^2$$

In order to work out the rate constant of this equation, the equation has to be re-arranged to:

$$k = \frac{\text{Rate}}{[\text{Sulfuric Acid}]^2}$$

By putting my experimental result from the 2.8 molar concentration series into the equation you get:

$$k = \frac{0.185}{2.8^2}$$

Therefore the rate constant 'k' is equal to 0.0236 to 3 significant figures.

I have identified a possible mechanism for the production of hydrogen from the uncatalysed reaction of zinc and sulfuric acid. This is illustrated in the following

diagram.

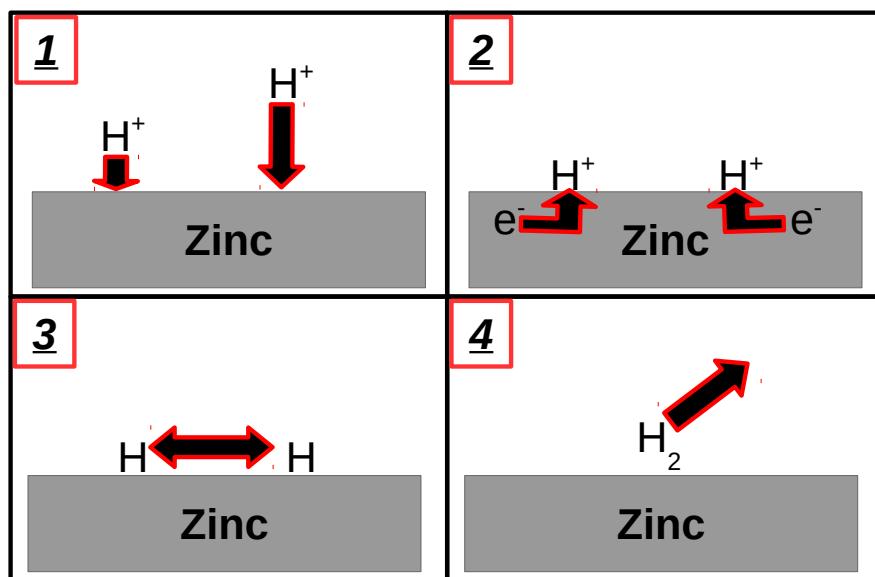


Figure 5.9: Proposed mechanism for the zinc acid reaction

1. Hydrogen ions supplied from the acid arrive at the powdered zinc's surface area.
2. Electrons supplied from the zinc reduce hydrogen the hydrogen ions on the zinc's surface area, thus oxidising the zinc.
3. Hydrogen atoms form a covalent bond with each other to form a molecule of hydrogen gas.
4. The hydrogen gas moves away from the zinc's surface area.

From the proposed mechanism, the rate determining step can be

5.2 Zinc and Sulfuric Acid With a Copper Sulfate Catalyst

5.2.1 Sulfuric Acid

Below are the graphs that I have used to work out the rate of the reaction for the experiment series that did involve a catalyst but varied in sulfuric acid concentration. All graphs have taken a mean average of the three repeats in order to plot the data points. As discussed in my 'Chosen Method' section on page 33 I am using the equation $\text{Gradient} = \frac{\text{Change in Y}}{\text{Change in X}}$ and as the gradient is equal to the rate of reaction, I will be able to use the graphs below in order to generate a rate vs concentration graph which will determine the order of the reaction. I will calculate all rates to 3 decimal places.

Graph To Show How Much Hydrogen Was Produced From 1.0 g of Zinc and 40 cm³ of 0.2 Molar Sulfuric Acid and 0.01 Molar Copper Sulfate.

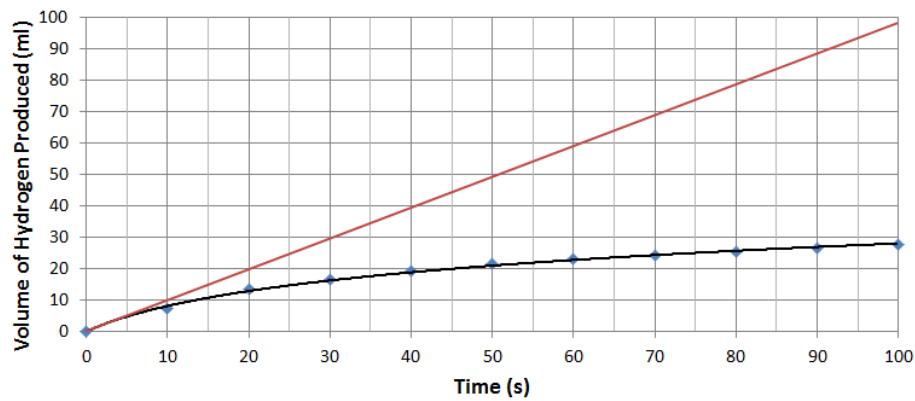


Figure 5.10: 0.2 Molar Sulfuric Acid and 0.01 Molar Copper Sulfate and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{97}{100}$$

Therefore:

$$\text{Rate} = 0.970 \text{ ml s}^{-1}$$

Graph To Show How Much Hydrogen Was Produced From 1.0 g of Zinc and 40 cm³ of 0.4 Molar Sulfuric Acid and 0.01 Molar Copper Sulfate.

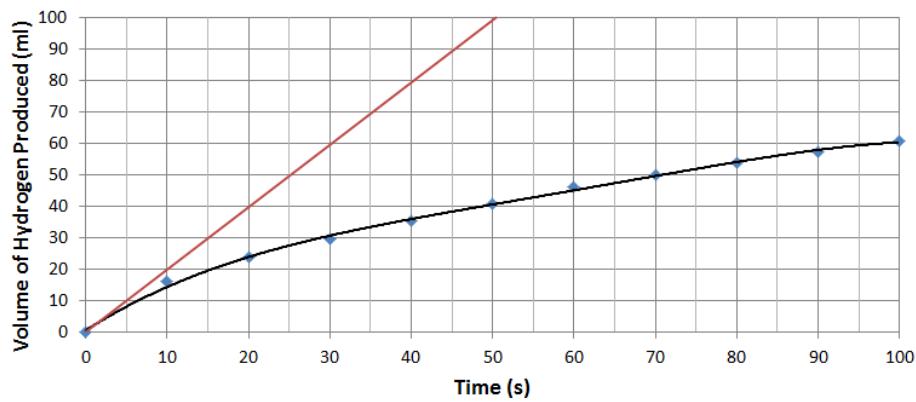


Figure 5.11: 0.4 Molar Sulfuric Acid and 0.01 Molar Copper Sulfate and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{100}{51.5}$$

Therefore:

$$\text{Rate} = 1.942 \text{ ml s}^{-1}$$

Graph To Show How Much Hydrogen Was Produced From 1.0 g of Zinc and 40 cm³ of 0.6 Molar Sulfuric Acid and 0.01 Molar Copper Sulfate.

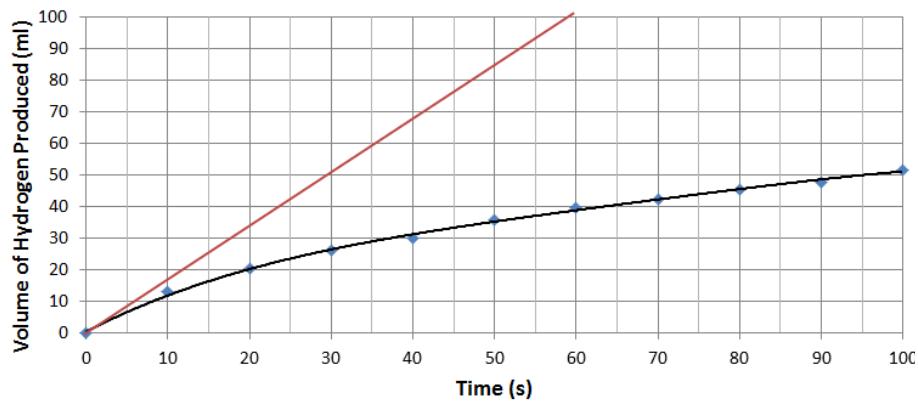


Figure 5.12: 0.6 Molar Sulfuric Acid and 0.01 Molar Copper Sulfate and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{100}{59}$$

Therefore:

$$\text{Rate} = 1.695 \text{ ml s}^{-1}$$

Graph To Show How Much Hydrogen Was Produced From 1.0 g of Zinc and 40 cm³ of 0.8 Molar Sulfuric Acid and 0.01 Molar Copper Sulfate.

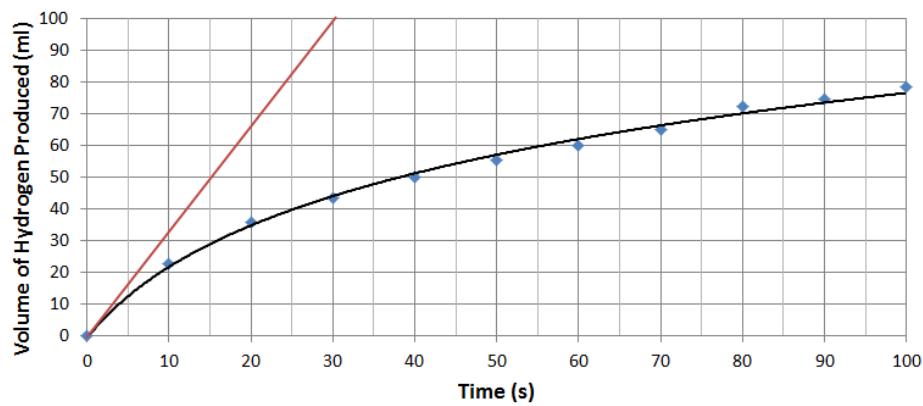


Figure 5.13: 0.8 Molar Sulfuric Acid and 0.01 Molar Copper Sulfate and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{100}{30.7}$$

Therefore:

$$\text{Rate} = 3.257 \text{ ml s}^{-1}$$

Graph To Show How Much Hydrogen Was Produced From 1.0 g of Zinc and 40 cm³ of 1.0 Molar Sulfuric Acid and 0.01 Molar Copper Sulfate.

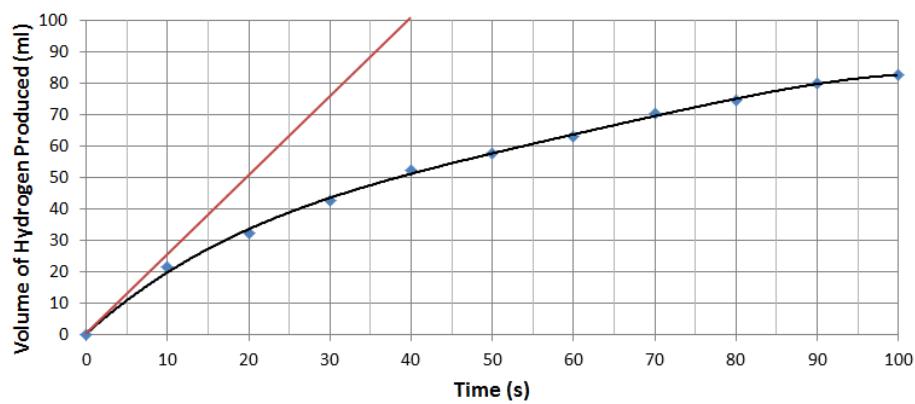


Figure 5.14: 1.0 Molar Sulfuric Acid and 0.01 Molar Copper Sulfate and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{100}{40}$$

Therefore:

$$\text{Rate} = 2.500 \text{ ml s}^{-1}$$

Graph To Show How Much Hydrogen Was Produced From 1.0 g of Zinc and 40 cm³ of 1.2 Molar Sulfuric Acid and 0.01 Molar Copper Sulfate.

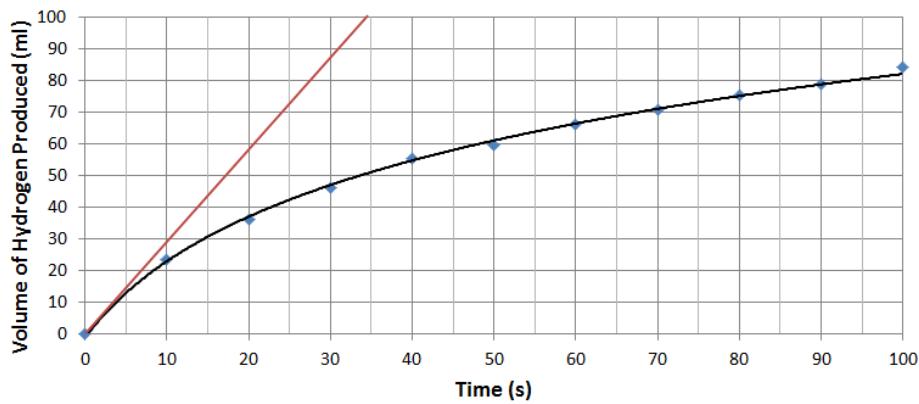


Figure 5.15: 1.2 Molar Sulfuric Acid and 0.01 Molar Copper Sulfate and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{100}{34.5}$$

Therefore:

$$\text{Rate} = 2.899 \text{ ml s}^{-1}$$

Below the rates of reactions are summarised in a table.

Sulfuric Acid Concentration (Molar)	Initial Rate of Reaction (ml s^{-1})
0.2	0.97
0.4	1.94
0.6	1.69
0.8	3.26
1.0	2.50
1.2	2.90

Figure 5.16: Initial Rates of Reactions for the Catalysed Zinc and Sulfuric Acid Reaction

Below the rates of reactions are compiled into a rate vs concentration graph.

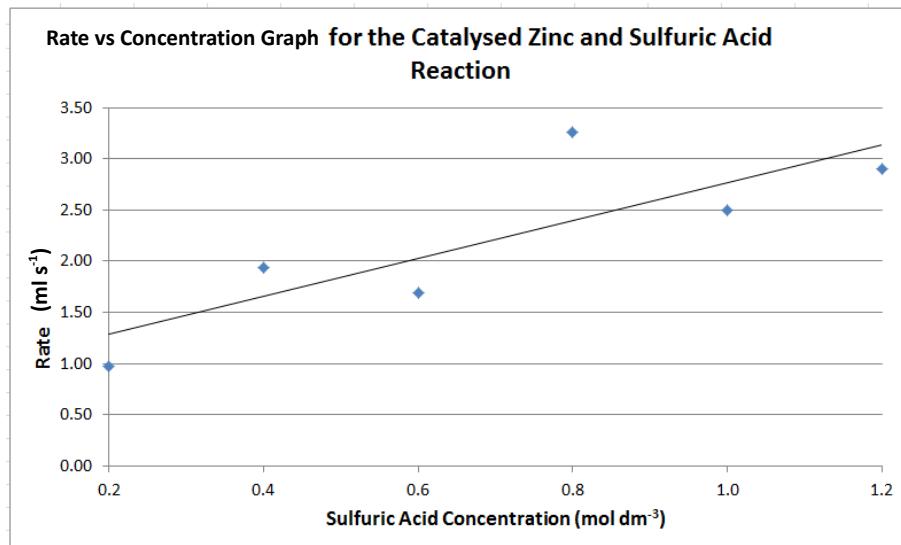


Figure 5.17: Rate vs Concentration Graph for the Catalysed Zinc and Sulfuric Acid Reaction

The rate vs concentration graph for this experiment is too scattered to get an accurate result; however due to the scattered result it appears that a straight

line of best fit would be most appropriate. Therefore I believe that the when in the presence of copper sulfate, sulfuric acid is a first order reactant. This means that doubling the concentration of sulfuric acid in this reaction doubles the initial rate of reaction. The rate equation for this reaction would be:

$$\text{Rate} = k[\text{Sulfuric Acid}]$$

Next I will work out if copper sulfate is also part of the rate equation.

5.2.2 Copper Sulfate

Below are the graphs that I have used to work out the rate of the reaction for the experiment series that did involve a catalyst and varied in copper sulfate concentration. All graphs have taken a mean average of the three repeats in order to plot the data points. As discussed in my 'Chosen Method' section on page 33 I am using the equation $\text{Gradient} = \frac{\text{Change in Y}}{\text{Change in X}}$ and as the gradient is equal to the rate of reaction, I will be able to use the graphs below in order to generate a rate vs concentration graph which will determine the order of the reaction. I will calculate all rates to 3 decimal places.

**Graph To Show The Volume Of Hydrogen Produced From
1.0 g of Zinc and 40 cm³ 0.2 Molar Sulfuric Acid and 0.01
Molar Copper Sulfate**

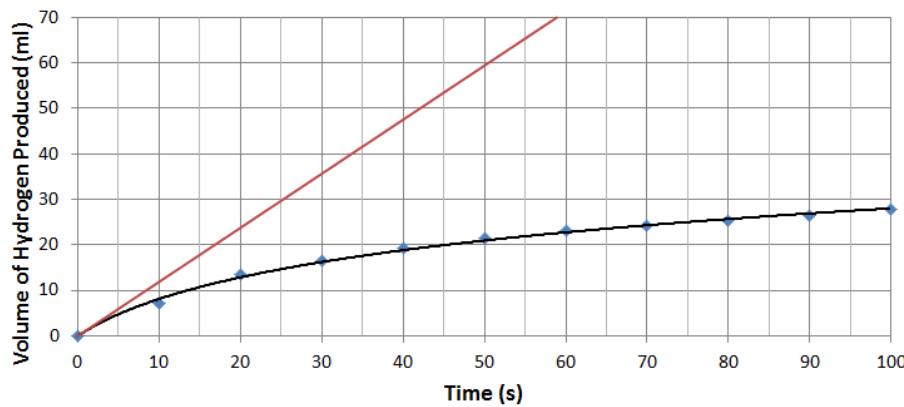


Figure 5.18: 0.4 Molar Sulfuric Acid and 0.01 Molar Copper Sulfate and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{70}{58}$$

Therefore:

$$\text{Rate} = 1.207 \text{ ml s}^{-1}$$

Graph To Show The Volume Of Hydrogen Produced From 1.0 g of Zinc and 40 cm³ 0.2 Molar Sulfuric Acid and 0.02 Molar Copper Sulfate

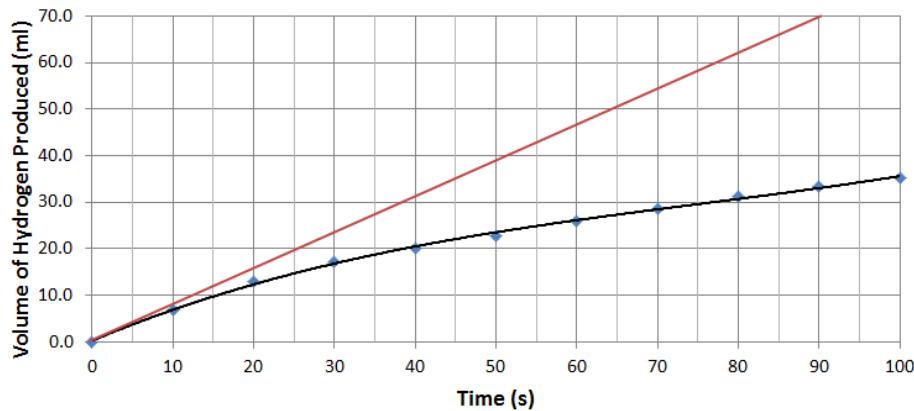


Figure 5.19: 0.4 Molar Sulfuric Acid and 0.02 Molar Copper Sulfate and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{79}{90}$$

Therefore:

$$\text{Rate} = 0.778 \text{ ml s}^{-1}$$

Graph To Show The Volume Of Hydrogen Produced From 1.0 g of Zinc and 40 cm³ 0.2 Molar Sulfuric Acid and 0.03 Molar Copper Sulfate

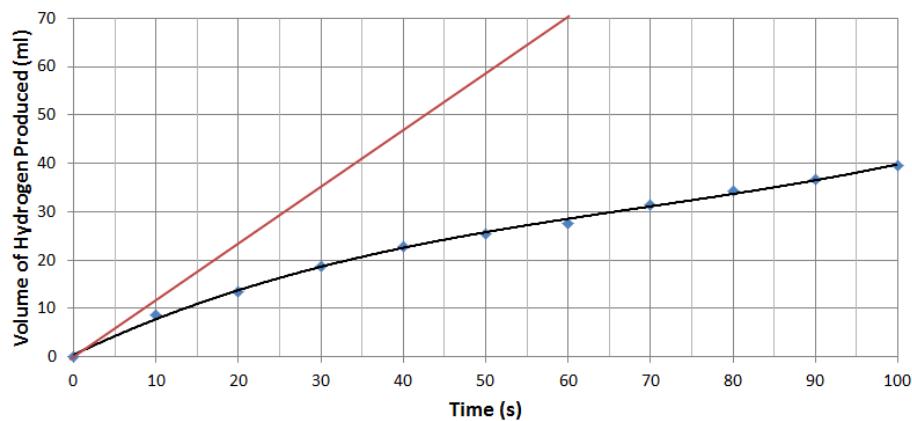


Figure 5.20: 0.4 Molar Sulfuric Acid and 0.03 Molar Copper Sulfate and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{70}{60}$$

Therefore:

$$\text{Rate} = 1.167 \text{ ml s}^{-1}$$

**Graph To Show The Volume Of Hydrogen Produced From
1.0 g of Zinc and 40 cm³ 0.2 Molar Sulfuric Acid and 0.04
Molar Copper Sulfate**

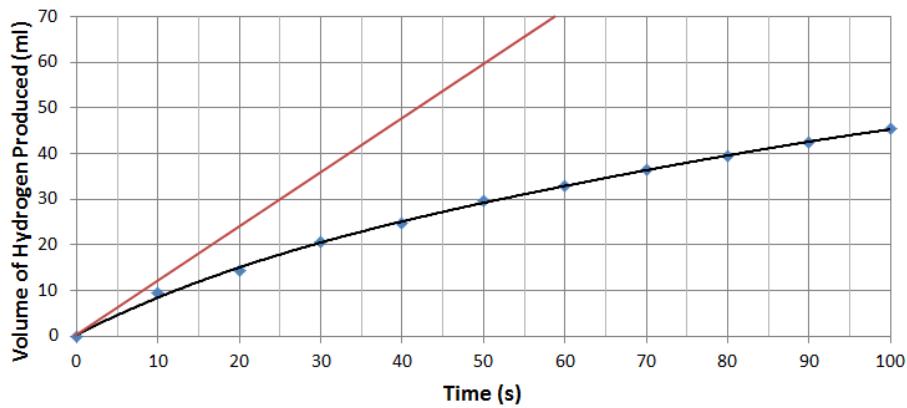


Figure 5.21: 0.4 Molar Sulfuric Acid and 0.04 Molar Copper Sulfate and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{70}{57.5}$$

Therefore:

$$\text{Rate} = 1.217 \text{ ml s}^{-1}$$

**Graph To Show The Volume Of Hydrogen Produced From 1.0 g
of Zinc and 40 cm³ 0.2 Molar Sulfuric Acid and 0.05 Molar
Copper Sulfate**

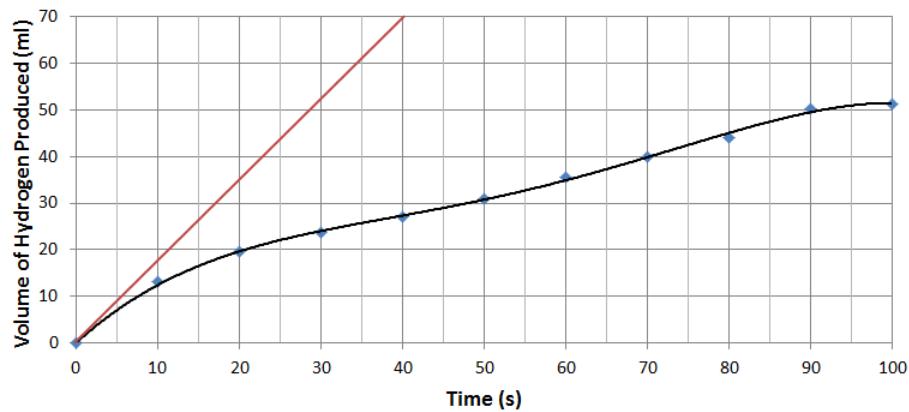


Figure 5.22: 0.4 Molar Sulfuric Acid and 0.05 Molar Copper Sulfate and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{70}{40}$$

Therefore:

$$\text{Rate} = 1.750 \text{ ml s}^{-1}$$

Graph To Show The Volume Of Hydrogen Produced From 1.0 g of Zinc and 40 cm³ 0.2 Molar Sulfuric Acid and 0.06 Molar Copper Sulfate

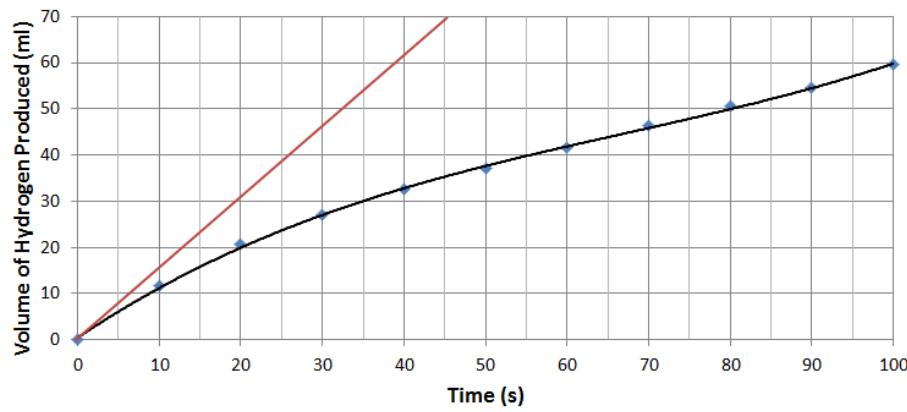


Figure 5.23: 0.4 Molar Sulfuric Acid and 0.06 Molar Copper Sulfate and 1.0 g of Zinc Average Graph

$$\text{Gradient} = \frac{70}{45}$$

Therefore:

$$\text{Rate} = 1.556 \text{ ml s}^{-1}$$

Below the rates of reactions are summarised in a table.

Concentration of Copper Sulfate (molar)	Initial Rate of Reaction (ml s ⁻¹)
0.01	1.207
0.02	0.778
0.03	1.167
0.04	1.217
0.05	1.750
0.06	1.556

Figure 5.24: Initial Rates of Reactions for the Catalysed Zinc and Sulfuric Acid Reaction

Below the rates of reactions are compiled into a rate vs concentration graph.

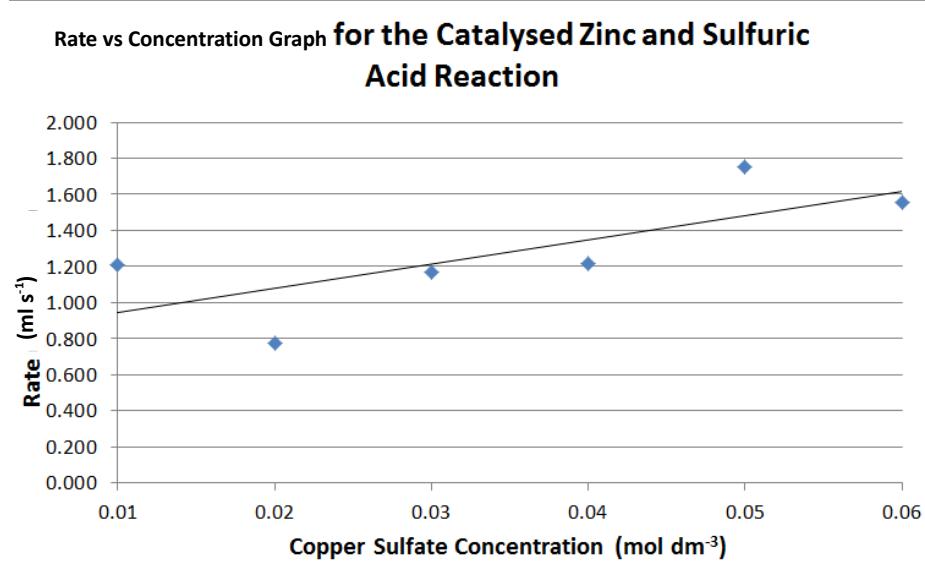


Figure 5.25: Rate vs Concentration Graph for the Catalysed Zinc and Sulfuric Acid Reaction

The rate vs concentration graph for this experiment, again, is too scattered to get an accurate result; however due to the scattered result it appears that a

straight line of best fit would be most appropriate. Therefore I believe that the copper sulfate is a first order reactant. This means that doubling the concentration of copper sulfate in this reaction doubles the initial rate of reaction. The rate equation for this overall reaction would be:

$$\text{Rate} = k[\text{Sulfuric Acid}][\text{Copper Sulfate}]$$

This means that overall the reaction is a second order reaction and both individual reactants are first order reactants.

5.3 Comparing Catalysts

The copper sulfate graph can be found on page 73, Figure 5.11.

Graph To Show The Volume Of Hydrogen Produced From 1 g of Zinc and 40 cm³ 0.4 Molar Sulfuric Acid and 0.01 Molar Nickel Sulfate

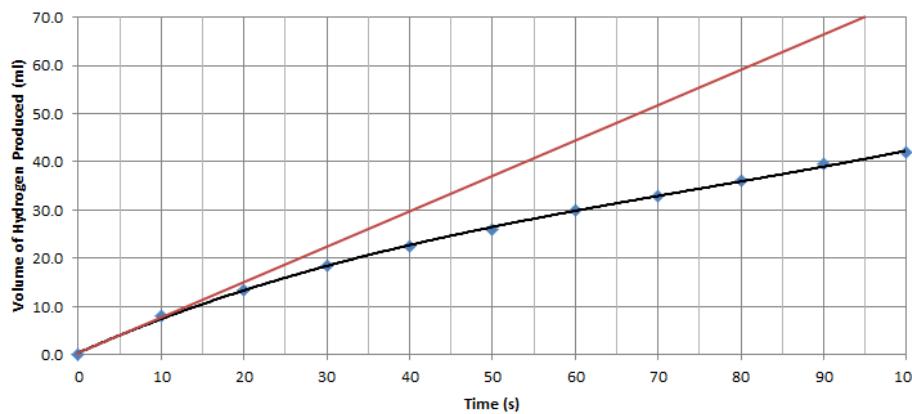


Figure 5.26: 0.4 Molar Sulfuric Acid and 0.01 Molar Nickel Sulfate and 1.0 g of Zinc Average Graph

Graph To Show The Volume Of Hydrogen Produced From 1 g of Zinc and 40 cm³ 0.4 Molar Sulfuric Acid and 0.01 Molar Manganese Sulfate

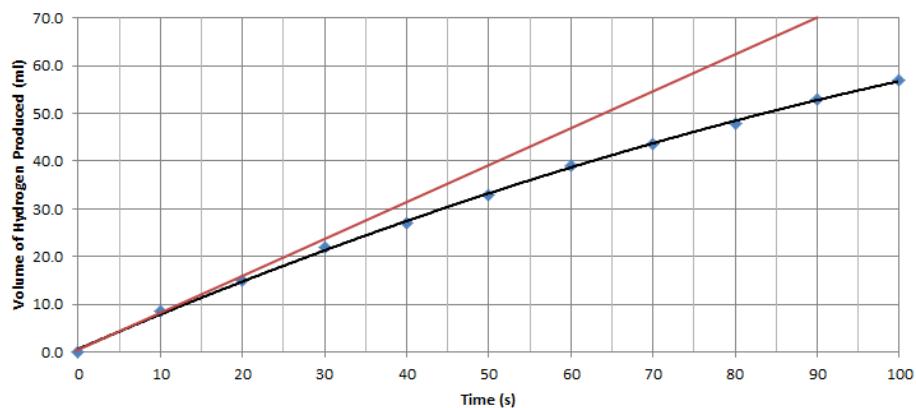


Figure 5.27: 0.4 Molar Sulfuric Acid and 0.01 Molar Manganese Sulfate and 1.0 g of Zinc Average Graph

Graph To Show The Volume Of Hydrogen Produced From 1 g of Zinc and 40 cm³ 0.4 Molar Sulfuric Acid and 0.01 Molar Iron Sulfate

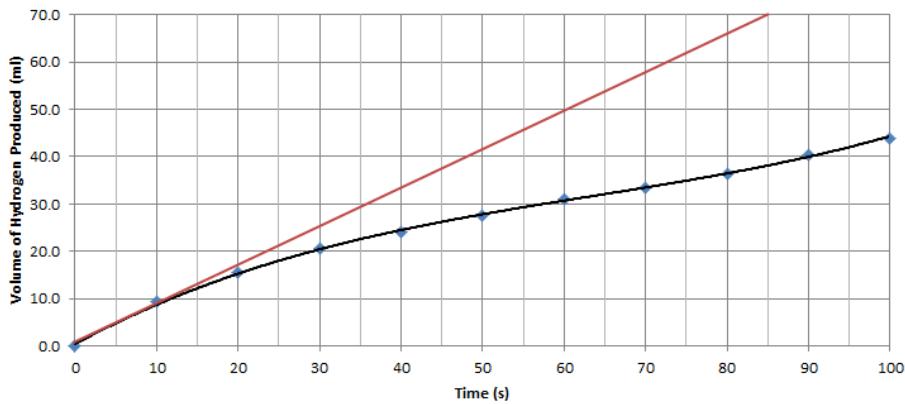


Figure 5.28: 0.4 Molar Sulfuric Acid and 0.01 Molar Iron Sulfate and 1.0 g of Zinc Average Graph

Below the rates of reactions are summarised in a table.

Metal	Initial Rate of Reaction (ml s ⁻¹)
Copper	1.842
Nickel	0.737
Manganese	0.778
Iron	0.824

Figure 5.29: Initial Rates of Reactions for the Different Catalysts Tested

For this experiment I have created a bar chart to compare the rate of reaction, this can be seen below.

Comparison of Catalysts Effectiveness to Catalyse the Zinc and Sulfuric Acid Reaction

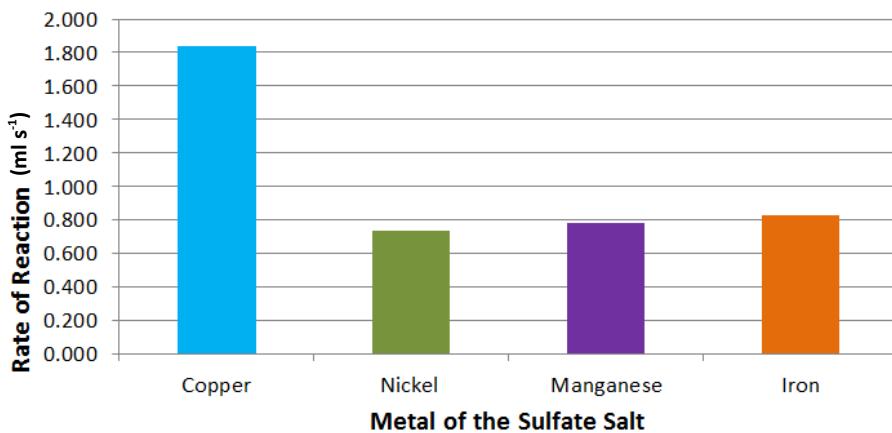


Figure 5.30: Comparison of Rates of Different Sulfate Catalysts

This experiment has shown that copper sulfate is the most effective catalyst, which has over double the rate of all the other sulfates. Iron is the next most effective catalyst followed by manganese and then nickel.

I believe that copper is the most effective catalyst as the electrolytic cell that it creates works so well together as the copper is not as reactive as zinc.

Catalysts which are of similar reactivity to zinc (iron and nickel) do not create an as efficient electrolytic cell as that of copper which is significantly less reactive than zinc. By having the catalyst significantly less reactive than zinc the electrolytic cell becomes more efficient at producing hydrogen as hydrogen ions are going to be reduced to form hydrogen gas (one of the products of a metal and acid reaction) at the less reactive electrode. This leads to an increased production of hydrogen gas. The manganese sulfate result can be explained by the fact that manganese is more reactive than zinc, and therefore does not cause an increase in production of hydrogen in the same way as the other catalysts tested. I think that this is the reason that copper sulfate is a significantly more efficient catalyst.

5.4 Investigating the Solubility of Hydrogen

The graph below shows all of the data points taken from the raw data table in the Experiment section of my coursework.

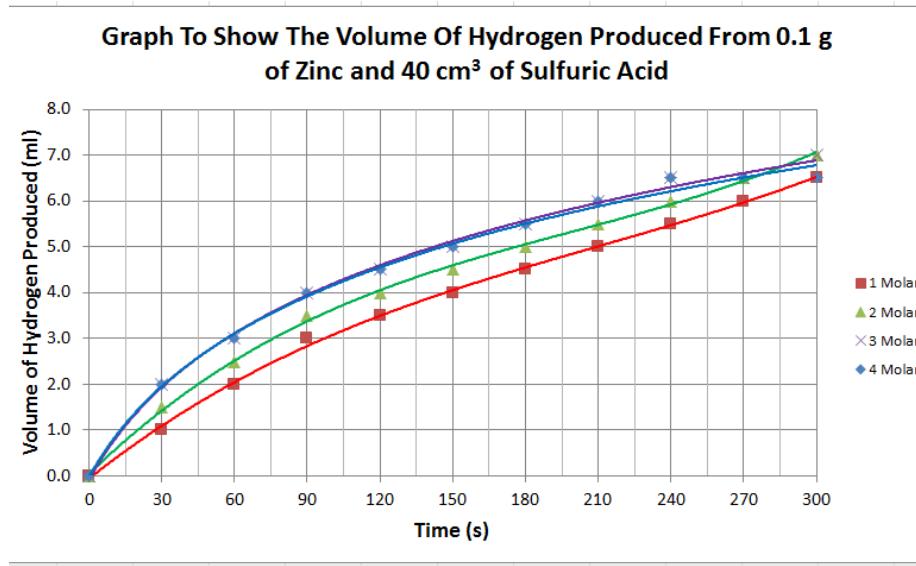


Figure 5.31: Comparison of the Hydrogen Produced When Different Concentrations of Sulfuric Acid is in Excess

Although there have been no scientific sources that I can find which discuss the solubility of hydrogen in higher concentrations of acids, my data starts to suggest that it does increase. At 4 molar sulfuric acid the volume of hydrogen produced is less than that of the lower concentrations which suggests that solubility of hydrogen does increase. Therefore the higher the concentration of the acid, the less accurate the results are as some of the hydrogen is dissolving into the solution. But to confidently comment on this, I would have to carry out the experiment with a wider range of acid concentrations and repeats.

From the equation of the reaction ($\text{Zn} (\text{s}) + \text{H}_2\text{SO}_4 (\text{aq}) \rightarrow \text{ZnSO}_4 (\text{aq}) + \text{H}_2 (\text{g})$) you can work out that that by using 40 cm³ of 4 molar sulfuric acid, the zinc is over 100x greater in excess. This calculation is shown below.

Moles of sulfuric acid:

- Number of Moles = Concentration × Volume
- $0.16 = 4 \times (40 \times 10^{-3})$

Moles of zinc:

- Number of Moles = $\frac{\text{Mass}}{\text{Molecular Mass}}$
- $0.001529285 = \frac{0.1}{65.39}$

Therefore there are over 100 more sulfuric acid particles than zinc. The total volume of hydrogen that can be produced by the reaction is 36.70 cm^3 . This can be worked out as the stoichiometry between zinc and hydrogen is a ratio of 1:1 and therefore the number of moles of hydrogen that can be produced is equal to the number of moles of the zinc (as the acid is in excess). This leaves the equation to be:

Volume of hydrogen produced:

- Volume = Number of Moles $\times 24$
- $0.03670284 \text{ dm}^3 = 0.001529285 \times 24$
- 36.70284 cm^3

Obviously my results has produced less hydrogen than expected for a number of reasons. First of all I only measured the volume of hydrogen produced over 5 minutes. Secondly, some hydrogen will have been produced before the bung was placed into the conical flask. Finally the zinc's surface area to volume ratio isn't are large as it could be, therefore not producing as much hydrogen as possible.

Chapter 6

Evaluation

6.1 Equipment

I used a volumetric pipette to get a specific measurement of a solution. The smallest division on the scale is 0.2 ml which leaves a +/- 0.1 ml uncertainty.

I used weighing scales to measure out a desired amount of zinc. The smallest division on the scale is 3 decimal places, therefore that leaves a +/- 0.0005 g uncertainty; however as I take two measurements (weighing the weighing boat and then the zinc) the uncertainty would double to 0.001 g.

I used a thermometer to measure the temperature of the reaction. As the smallest division on the scale is 1 degrees, the uncertainty would be 0.5 degrees. As my measurements were from 25 degrees - 26 degrees, the uncertainty would be in between the range of 24.5 degrees - 26.5 degrees.

I used a gas syringe which has the smallest division of 1 ml, therefore that leaves a +/- 0.5 ml uncertainty.

I used a 100cm³ volumetric flask. This has an uncertainty of 0.16 ml. Therefore the percentage uncertainty of this piece of equipment is worked out mathematically below.

$$\% \text{ uncertainty} = \frac{0.16 \times 1}{100} \times 100$$

$$\% \text{ uncertainty} = 0.16\%$$

I used a 200 cm³ volumetric flask. This has an uncertainty of 0.20 ml. Therefore the percentage uncertainty of this piece of equipment is worked out mathematically below.

$$\% \text{ uncertainty} = \frac{0.20 \times 1}{200} \times 100$$

$$\% \text{ uncertainty} = 0.10\%$$

I used a timer on my phone which measures every millisecond, the uncertainty of that is +/- 500 microseconds.

6.2 Problems

I have had a number of problems which will have effected my experiment. These are discussed below.

1. Stock Solutions - As I was relying on the chemicals being given to me to be the correct concentration, I had no control, or knowledge of any incorrectly made solutions. This could effect my measurements either way as I was purely relying on being given the correct concentrations I asked for. Next time I would ensure that the stock solutions that I was using were freshly ordered in.
2. Zinc Clumping - I found that as the concentration of sulfuric acid increased, the tendency for the zinc to form clumps increased. This meant that a reduced surface area was available for a reaction to take place and therefore less hydrogen was being produced than what should've been. This meant that at higher concentrations the rate of reaction seems to be lower than it actually is. If I were to carry out this experiment again, I would use a magnetic stirrer in order to reduce the zinc powder from clumping.
3. Temperature - As I didn't control the temperature of the reaction, the reaction will have heated up which in itself as the reaction is exothermic. This will therefore cause an increase in the rate of reaction. Additionally the hotter a specific day was, the faster the reaction seemed. If I were to carry out this experiment again I would carry the experiment out in a water bath so that I would control the experiments temperature.
4. Zinc Storage - As the zinc may have been stored in storage for a few years, zinc oxide may have formed around the surface area of the zinc powder. This means that some experiments that I carried out will have appeared to be slower than they should have been due to hydrogen gas not being produced. The equation for the reaction would be $ZnO + H_2SO_4 \rightarrow ZnSO_4 + H_2O$ which, as it produces water, would dilute the acid more which would make the reaction even more slow. If I were to carry out this experiment again I would ensure that the zinc that I was using was freshly ordered in and not stored for a long period of time.
5. Bung Problems - When the bung is pushed into the conical flask, the pressure of the system increases which causes the gas syringe to falsely start moving. This causes a few ml of gas to be falsely collected and therefore over estimates the volume of hydrogen produced. If I were to carry out this experiment again I would use a gas tap so that there was no

change in the pressure of the system when you close it. Another problem with the bung is the time it takes from adding the acid to the conical flask, to then placing the bung in the conical flask. Because of this, hydrogen gas is not collected straight away and therefore the volume of hydrogen collected is underestimated. If I were to carry this experiment out again I would have a chamber in which reactants can be added whilst the system is still closed.

6. Timer Problems - As I did not instantly start the timer from the point of the reaction starting (latency between adding the acid to the conical flask, placing the bung in and then starting the timer) some of the initial production of hydrogen is not recorded. To combat this issue a closed system would have to be introduced in order to get all of the hydrogen produced from the absolute start of the reaction. Therefore my results underestimated the volume of hydrogen produced.
7. Hydrogen Travelling - Before the hydrogen gas starts pushing the gas syringe, the hydrogen gas has to pass through the tubes leading to the gas syringe. This means that the gas syringe doesn't precisely measure the amount of hydrogen produced and therefore my results underestimated the volume of hydrogen produced. If I were to carry out this experiment again I would reduce the distance between the reaction mixture and the gas syringe.
8. Cleaning Glassware Between Experiments - Between experiments whilst I am in the lab, as I did not have time to completely and thoroughly wash the conical flask (only cleaned with distilled water between experiments) non-visible residue left over from the last experiment series could still be present in the conical flask. This would act as an impurity and skew the results of my experiments. For example, if there was zinc residue more hydrogen gas would be produced than expected due to more zinc being available to react. To combat this issue I would use cleaning solution in addition to distilled water.

Chapter 7

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7.2 Figures

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