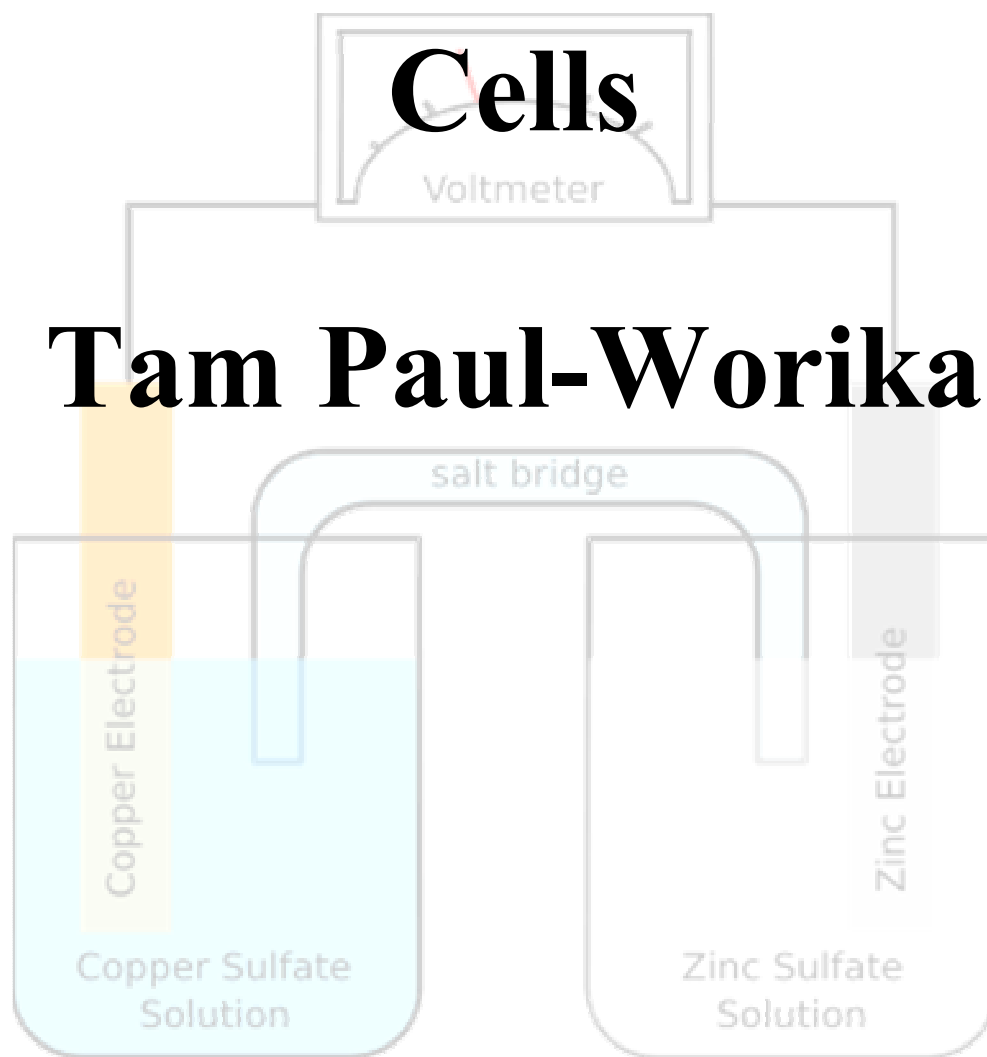


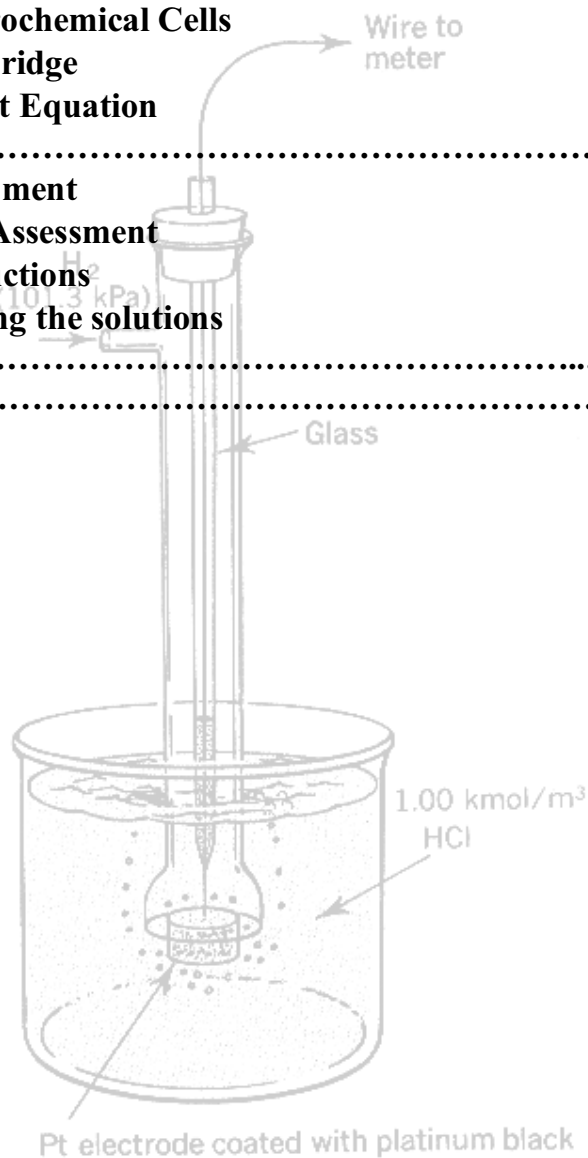
Electrochemical Cells

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Aim of the Investigation

In my investigation, I will observe how changing the concentration of a solution can affect the electrode potential in an electrochemical cell. I will also see if there is a trend when the concentration is changed.

Background

Redox Reactions

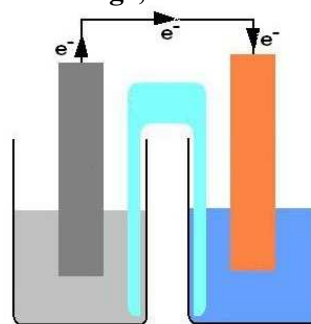
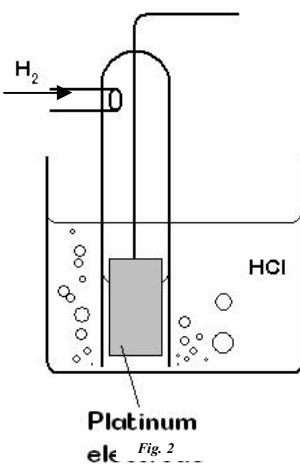
Redox reactions are reactions where the oxidation states of the atoms change^[1]. The atoms are either oxidised or reduced, depending on if they lose or gain electrons.

Electrochemical Cells

Electrochemical cells (also known as Galvanic cells or Voltaic cells^[2]), are devices that cause a current from chemical reactions. These reactions occur through redox.

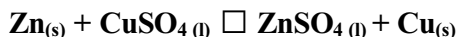
An electrochemical cell is set up so that the electrons lost from one of the reagents can travel to another reagent. This creates a voltage, which is also known as the electric potential difference. This voltage can be read if a high-resistance voltmeter is put into the circuit.

Electrochemical cells are made from two half-cells, which have their own standard electrode potential (E_0), which is measured in volts. This is measured by placing it with a standard hydrogen electrode. This is because for the half reaction that turns hydrogen ions into hydrogen gas, the standard electrode potential is 0.00V.



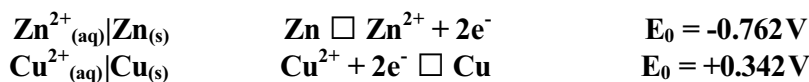
If a half-cell has a positive standard electrode potential, then it means that it is more likely to receive than give away electrons, which would make it the reducing agent. The oxidising agent is the negative half-cell.

For my investigation, I am using copper sulphate and zinc sulphate. When the solutions are 1 mol dm^{-3} , the redox reaction between them is:

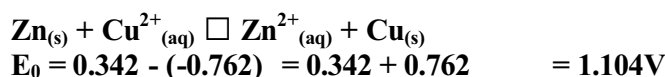


The redox reaction that occurs is a displacement reaction, as it would have been carried out in one reaction vessel. The reaction would release energy as heat. However, if the two reactions were split into two half-cells, that energy can be channelled through the flow of the electrons into electricity^[1].

The half-equations for the reactions are:



The standard electrode potential of the whole reaction is worked out by taking away the standard electrode potential of the anode, which will be the zinc, from the cathode (the copper). This means that for the ionic equation:



Salt Bridge^[3]

To complete the circuit of the cell, a salt bridge is used. The salt bridge is used to let the electrons flow from one of the solutions to another. For my investigation, I will be using filter paper soaked in saturated potassium nitrate (KNO_3). Potassium nitrate is used because it conducts electricity, but it will not react with either of the reagents.

Nernst Equation

Another way of working out the electrode potential is by using the Nernst equation^[4]. This method is mainly used when the electrochemical cell is not under standard conditions, such as having different concentrations. This is a simpler form of it:

$$E = E_0 - (0.05915/z) \times \log([red]/[ox])$$

- E is the electrode potential.
- E_0 is the standard electrode potential.
- z is the number of electrons transferred in the ionic equation.
- $[red]$ and $[ox]$ are the concentrations of the reduced and oxidated solutions respectively.





Plan

Equipment

- 10 x 100ml Beakers

- 2 x 50ml Measuring Cylinders
- 3 x 250ml Volumetric Flasks
- High-resistance Voltmeter
- 2 x Cables & Crocodile Clips
- Strips of Copper & Zinc
- Emery Paper
- Filter Paper
- Copper (II) Sulphate, 1 mol dm⁻³
- Zinc Sulphate, 1mol dm⁻³
- Potassium Nitrate, saturated
- Thermometer
- Scales

Risk Assessment^[5]

Substance	Hazard	Description	Precautions/Disposal
Copper (II) Sulphate		Harmful at 1M. Irritating to eyes and skin. Can sensitise the skin.	Wear gloves and goggles.
Zinc Sulphate		Toxic to the aquatic environment. May cause long-term adverse effects.	Copper (II) Sulphate: Dissolve in at least 8 litres of water for 40g before disposal. Zinc Sulphate: Dissolve in at least 4.05 litres of water for 40.5g before disposal.
Potassium Nitrate		Oxidising Agent. Contact with a combustible material may cause fire.	For 100g, dissolve in 10 litres of water before disposal.
Zinc Dust (from cleaning the zinc strips)		Flammable. Contact with water releases hydrogen gas	Destroy by adding slowly, with stirring, to excess dilute acid

Instructions

1. Make a 1mol dm⁻³ solution of copper (II) sulphate and zinc sulphate.
2. Pour 50cm³ of each into 100cm³ beakers. Using a folded piece of filter paper soaked in saturated potassium nitrate for the salt bridge, and a voltmeter (as shown in *fig. 3*), measure the electrode potential between the two solutions. Measure the temperature of the solutions at the time of the reaction.
3. Repeat number 2 for all of the different concentrations (1, 0.5, 0.25, 0.125, 0.625mol dm⁻³)
4. Keeping the copper (II) sulphate at 1mol dm⁻³, change the concentration of the zinc sulphate

Making the Solutions

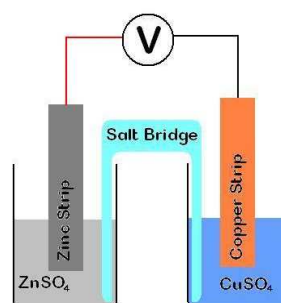


Fig. 3

All of the solutions that I am going to use in my investigation have to be made from the solid. Each of the compounds have different molar masses, so I have to use equations to work out how to make 1mol dm^{-3} of them in 250ml of water.

Copper (II) Sulphate: 1mol dm^{-3} for 250ml
Moles = $1 \times 250 \times 10^{-3} = 0.25\text{mol}$
Molar Mass = $64 + 32 + (4 \times 16) = 160\text{g/mol}$
Mass = $0.25 \times 160 = 40\text{g}$

Zinc Sulphate: 1mol dm^{-3} for 250ml
Moles = $1 \times 250 \times 10^{-3} = 0.25\text{mol}$
Molar Mass = $65 + 32 + (4 \times 16) = 161\text{g/mol}$
Mass = $0.25 \times 161 = 40.25\text{g}$

I was originally going to use a maximum of 2mol dm^{-3} for the solutions, but the maximum solubility for copper (II) sulphate is 79g for 250ml^[6], which means that the concentration would not be exactly 2mol dm^{-3} , as I would need 80g of the solid. Making the saturated potassium nitrate was easier to do because I needed at least 36g per 100ml^[7].

To make a dilute concentration from the original solution, I will have to use a ratio of water to the solution.

Concentration (mol dm^{-3})	Amount of water (%)	Amount of solution (%)
1	0	100
0.5	50	50
0.25	75	25
0.125	87.5	12.5
0.0625	93.75	6.25

This means that if I want to make 50ml of CuSO_4 at 0.25mol dm^{-3} , I would need to add 12.5ml of CuSO_4 to 37.5ml of water.

Analysis

Table to show Electrode Potentials of Zinc and Copper Ions (V)					
	Concentration of Zinc Sulphate (mol dm^{-3})				
	1.0	0.5	0.25	0.125	0.0625

Concentration of Copper (II) Sulphate (mol dm⁻³)	1.0	1.104	1.084	1.063	1.043	1.022
	0.5	1.124	1.104	1.084	1.063	1.043
	0.25	1.145	1.124	1.104	1.084	1.063
	0.125	1.165	1.145	1.124	1.104	1.084
	0.0625	1.186	1.165	1.145	1.124	1.104

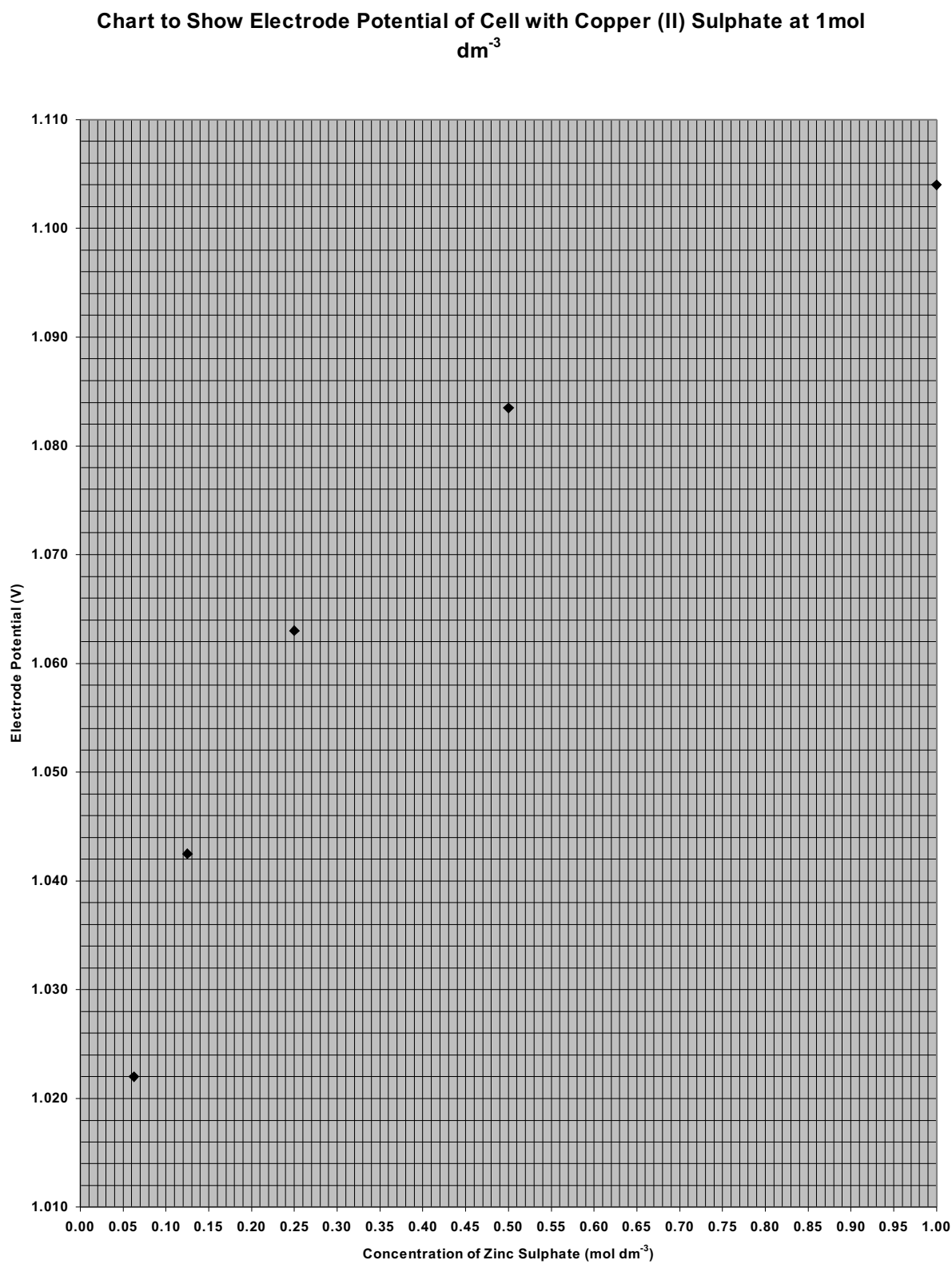


Chart to show Electrode Potential of Cell with Copper (II) Sulphate at 0.5mol dm^{-3}

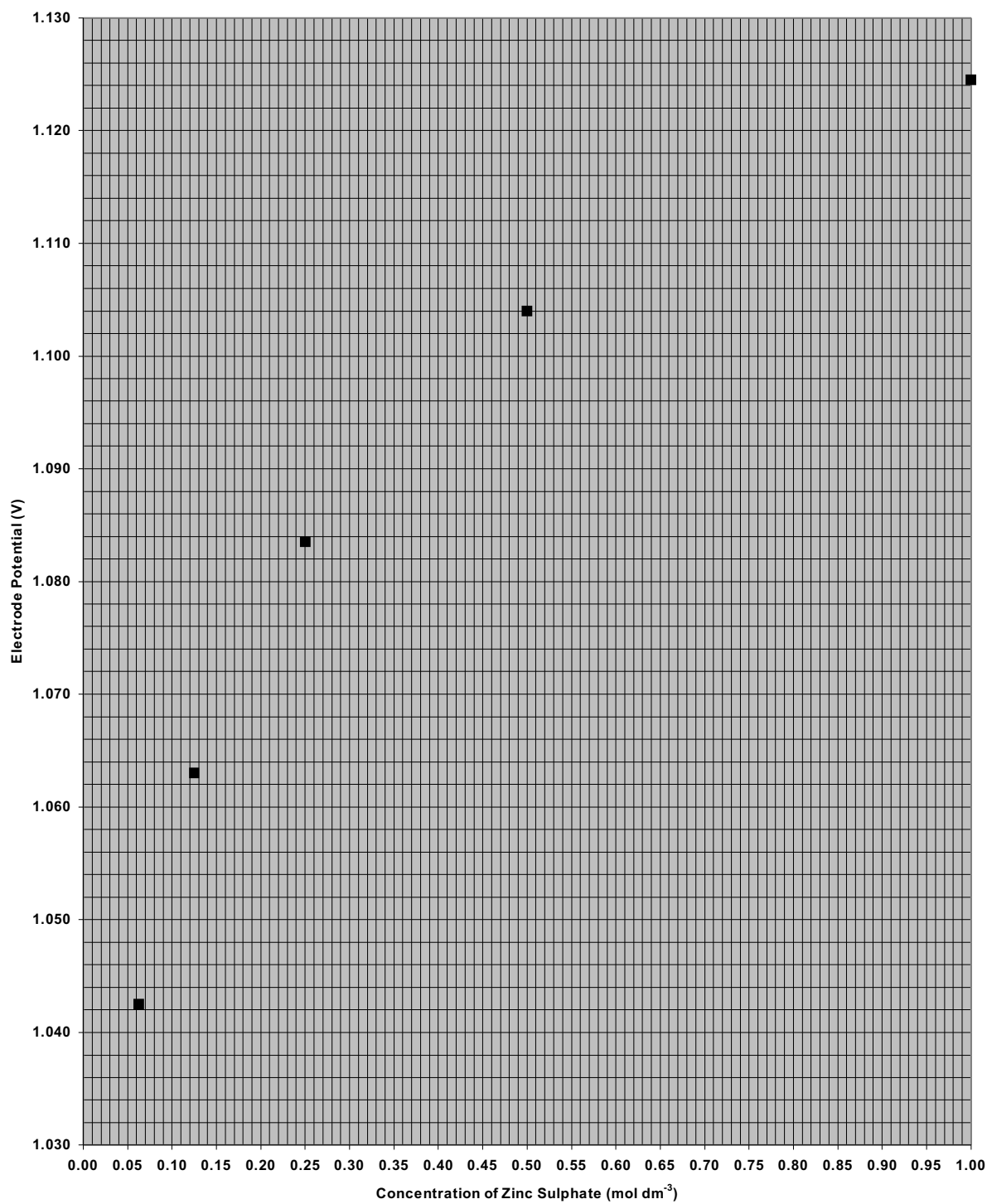


Chart to Show Electrode Potential of Cell with Copper (II) Sulphate at 0.25mol dm^{-3}

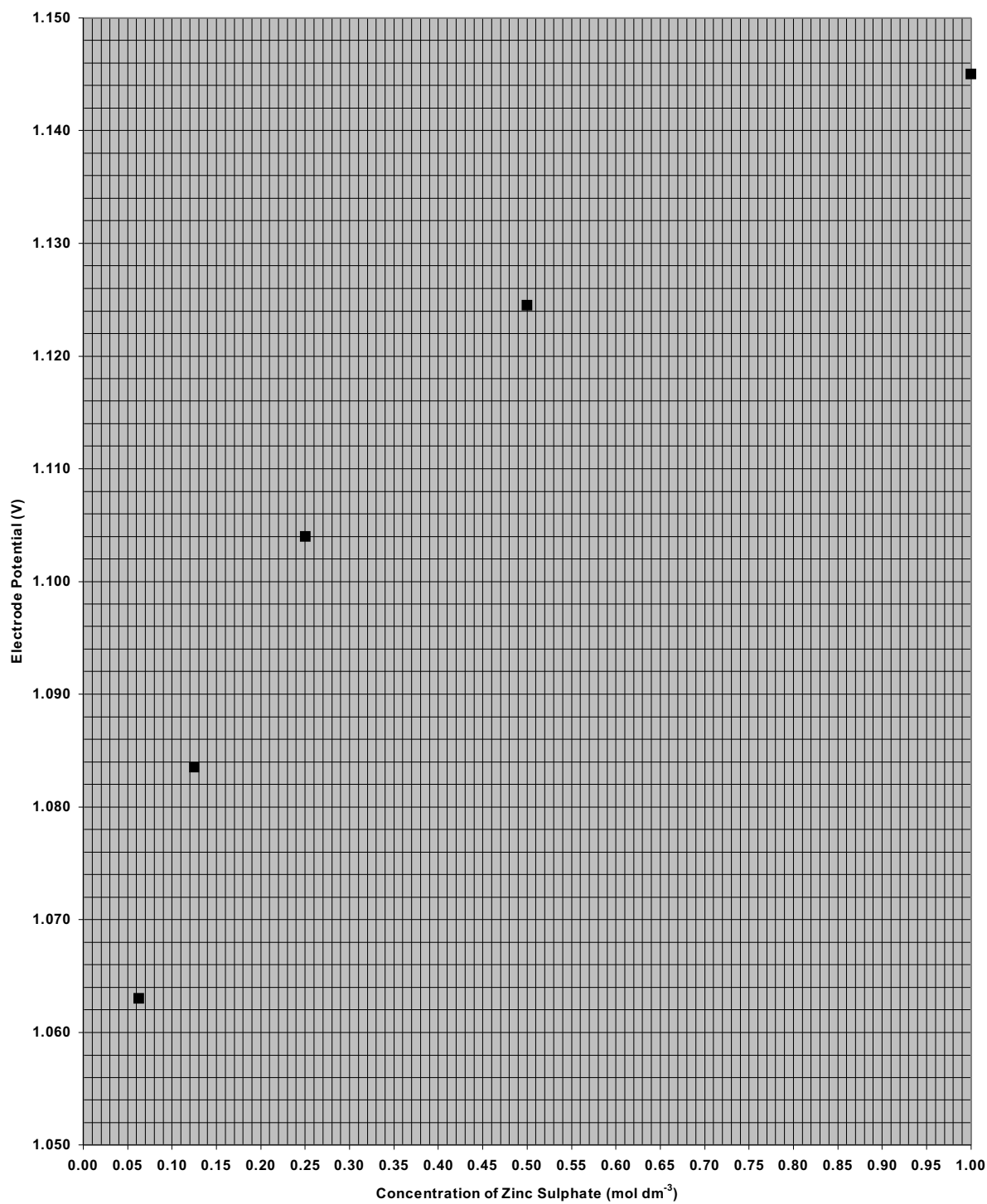


Chart to show Electrode Potential Cell with Copper (II) Sulphate at $0.125 \text{ mol dm}^{-3}$

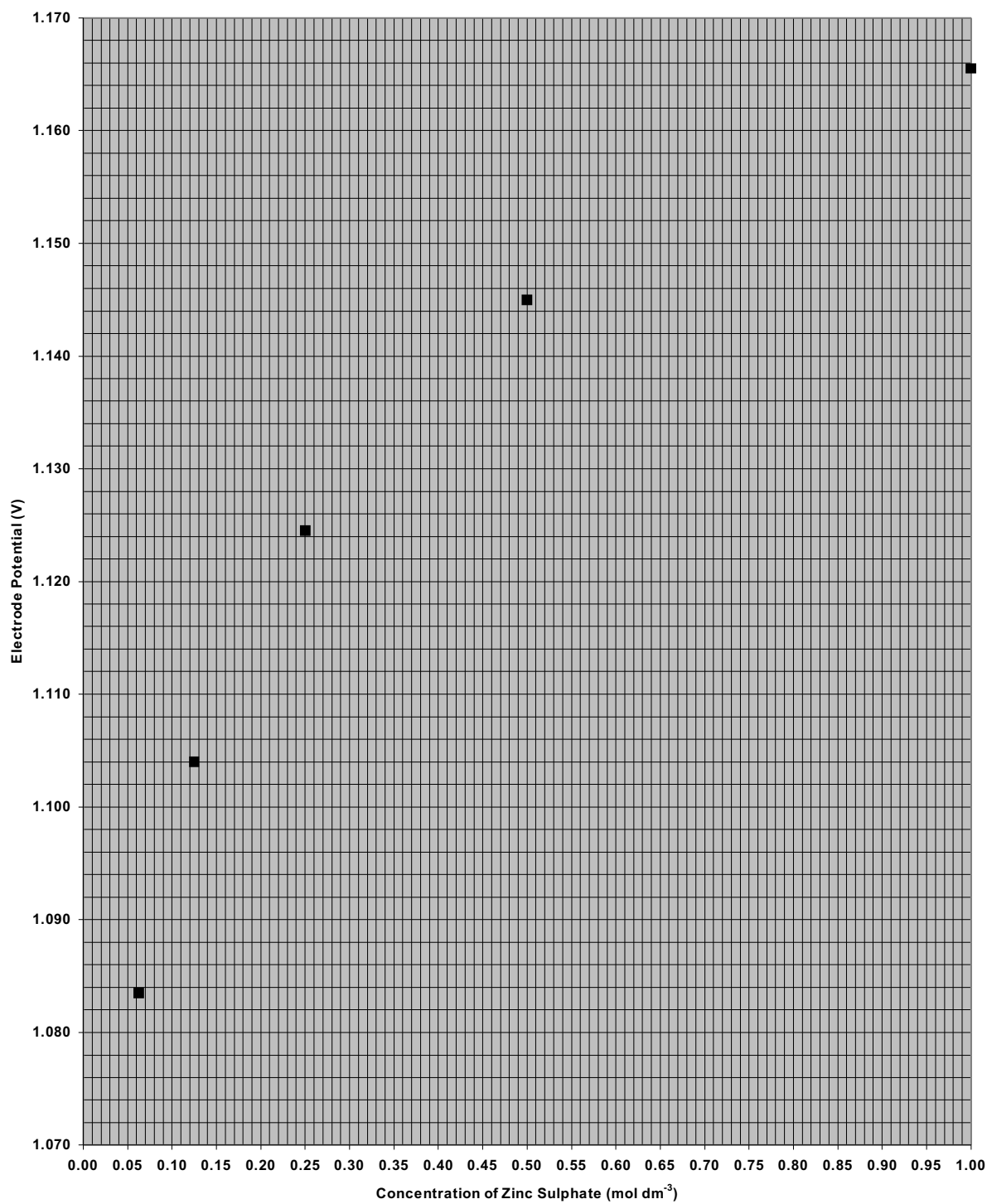


Chart to show Electrode Potential of Cell with Copper (II) Sulphate at $0.0625 \text{ mol dm}^{-3}$

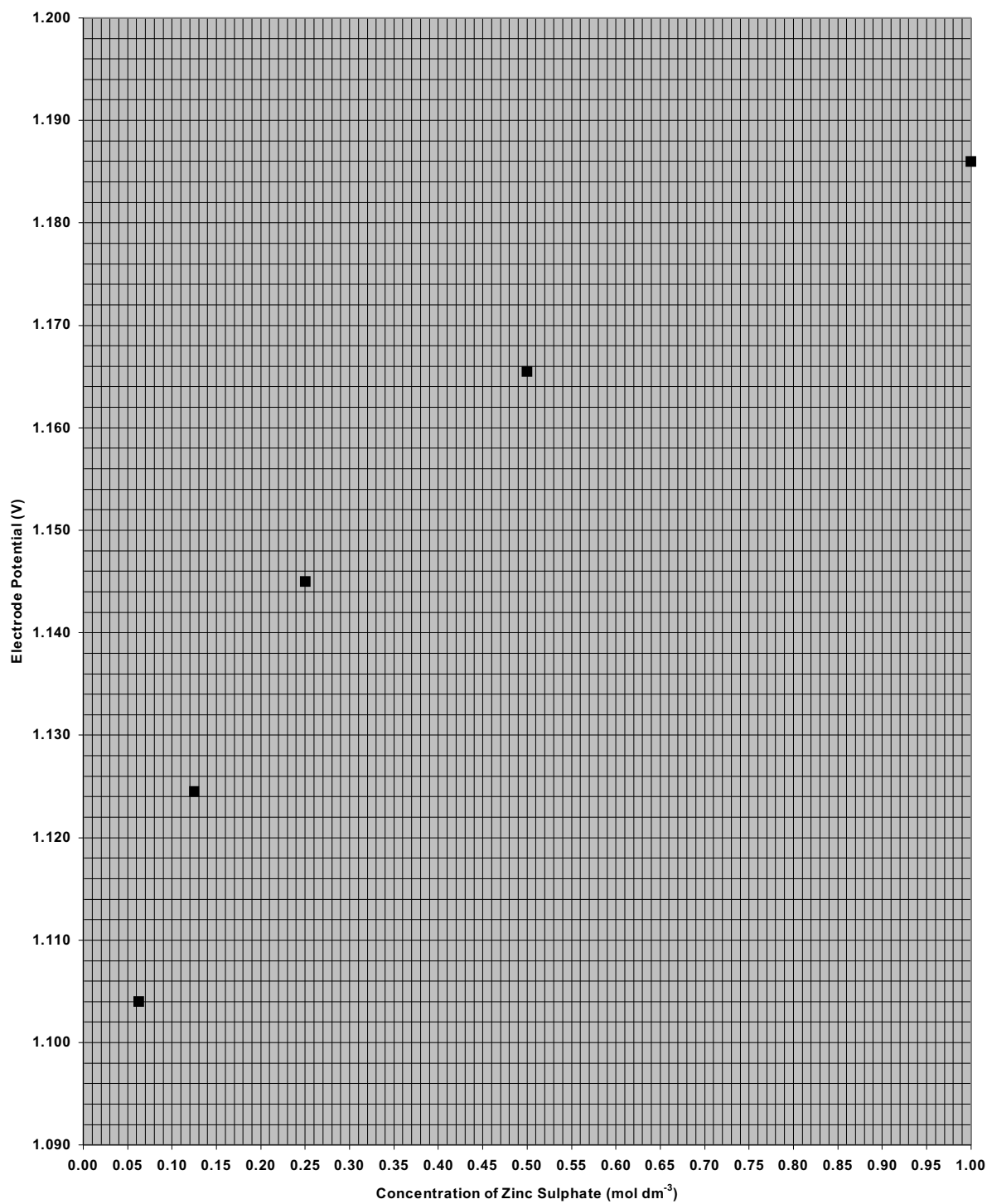


Chart to Show Electrode Potentials of Cells

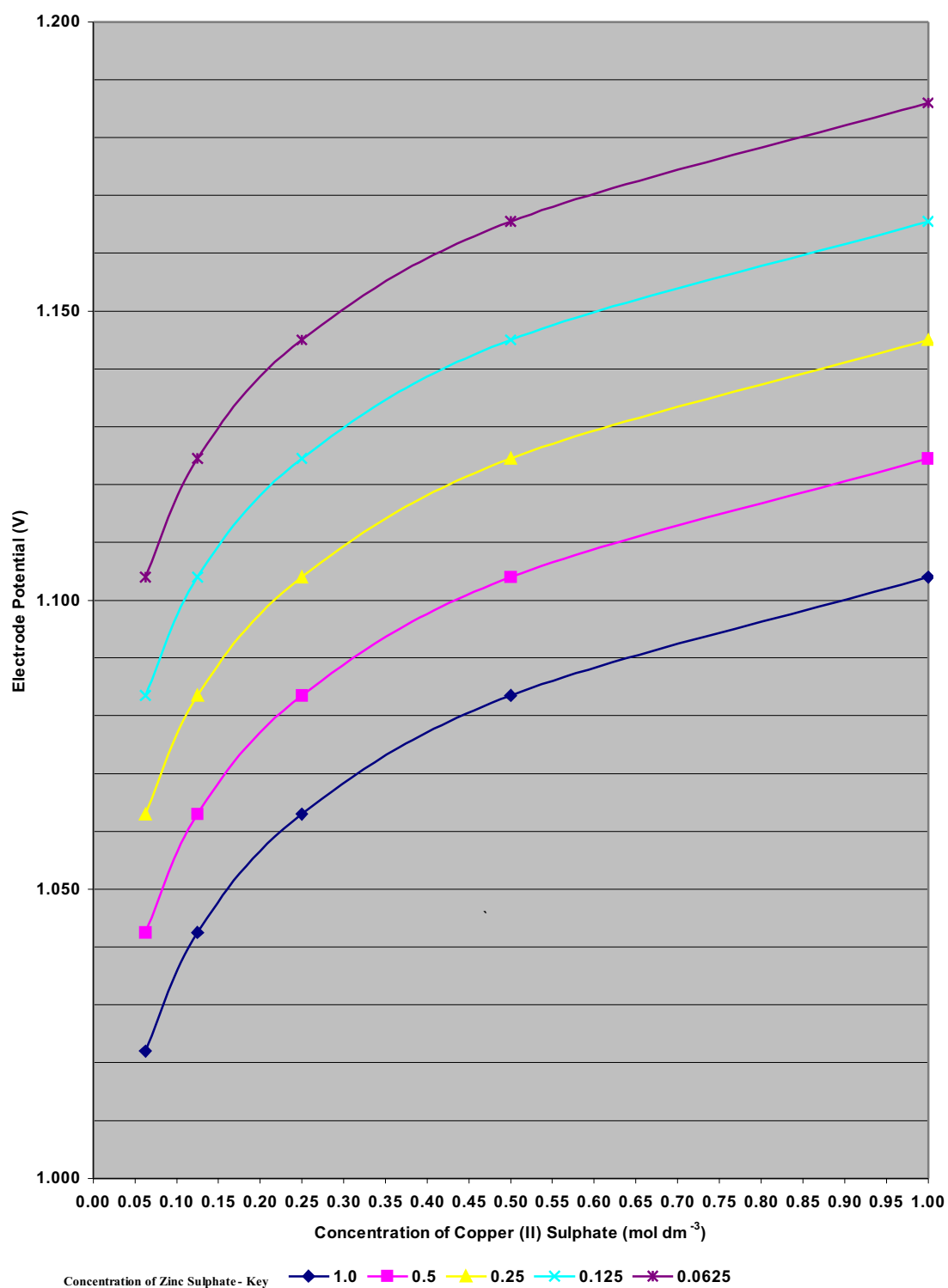
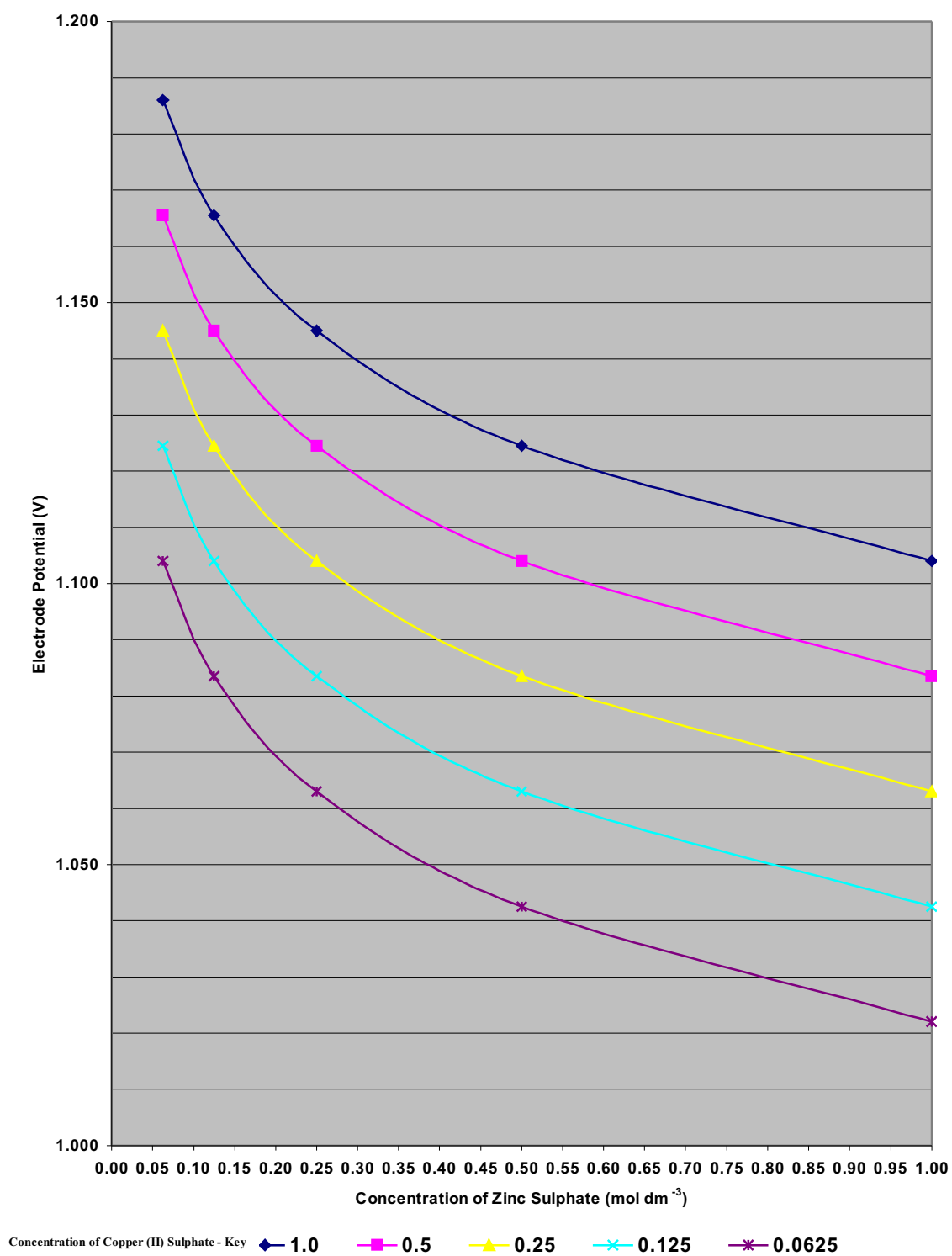
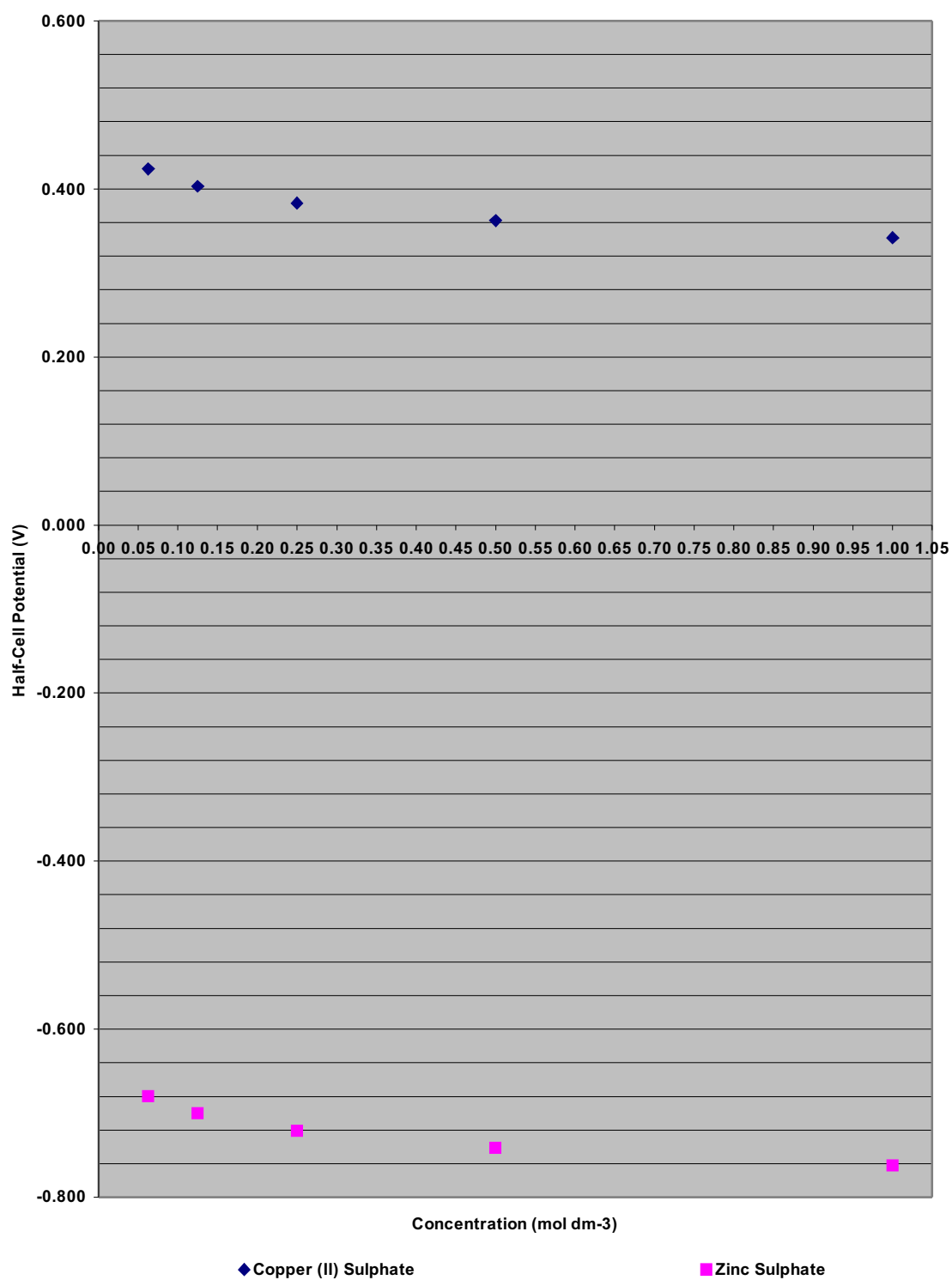


Chart to Show Electrode Potential of Cells



Graph showing Half-Cell Potentials				
Concentration of Copper Sulphate (mol dm ⁻³)	Concentration of Zinc Sulphate (mol dm ⁻³)	Electrode Potential (V)	Half-Cell Potential (V)	
			Copper (II) Sulphate	Zinc Sulphate
1	1	1.104	0.342	-0.762
	0.5	1.084	0.342	-0.742
	0.25	1.063	0.342	-0.721
	0.125	1.043	0.342	-0.701
	0.0625	1.022	0.342	-0.680
0.5	1	1.124	0.362	-0.762
	0.5	1.104	0.362	-0.742
	0.25	1.084	0.362	-0.721
	0.125	1.063	0.362	-0.701
	0.0625	1.043	0.362	-0.680
0.25	1	1.145	0.383	-0.762
	0.5	1.124	0.383	-0.742
	0.25	1.104	0.383	-0.721
	0.125	1.084	0.383	-0.701
	0.0625	1.063	0.383	-0.680
0.125	1	1.165	0.403	-0.762
	0.5	1.145	0.403	-0.742
	0.25	1.124	0.403	-0.721
	0.125	1.104	0.403	-0.701
	0.0625	1.084	0.403	-0.680
0.0625	1	1.186	0.424	-0.762
	0.5	1.165	0.424	-0.742
	0.25	1.145	0.424	-0.721
	0.125	1.124	0.424	-0.701
	0.0625	1.104	0.424	-0.680

Chart to show Half-Cell Potentials



From the data, you can see that there is a connection between the concentration and the electrode potential. As the concentration of the copper (II) sulphate increases, the electrode potential increases. On the other hand, when the concentration of the zinc sulphate is increased, the electrode potential decreases.

You can also see that the electrode potential increases when the concentration of copper (II) sulphate decreases. The half-cell potential of copper (II) sulphate also increases when the concentration decreases. This happens because as the concentration decreases, the number of electrons in the solution is less than usual. Because of this, the solution is even more likely to receive more electrons, making it more positive. The half-cell potential of zinc sulphate also increases because it is less likely to give away electrons than usual.

In the graphs, you can see that when you compare the electrode potentials for each concentration of copper (II) sulphate, all of the lines are almost parallel to each other. This shows that if, for example, you have 1mol dm⁻³ of copper (II) sulphate and 0.5mol dm⁻³ of zinc sulphate, you can tell that the electrode potential is 1.084V. If you change the concentration of copper (II) sulphate to 0.5mol dm⁻³, you would have to add 0.020V to it, and then if you had 0.25mol dm⁻³ of copper (II) sulphate, you add 0.021V. For each one, you add around 0.020V, so you can work out the approximate electrode potential.

Evaluation

Measurement Errors

Beakers:

$$\begin{aligned} \frac{0.05\text{ml}}{50\text{ml}} \times 100 &= 0.1\% \\ \frac{0.05\text{ml}}{200\text{ml}} \times 100 &= 0.025\% \\ \frac{0.05\text{ml}}{250\text{ml}} \times 100 &= 0.02\% \end{aligned}$$

Measuring Cylinders:

$$\frac{0.05\text{ml}}{50\text{ml}} \times 100 = 0.1\%$$

Thermometer:

$$\pm 0.5^{\circ}\text{C}$$

Voltmeter:

$$\pm 0.0005\text{V}$$

Volumetric Flask:

$$\frac{0.5\text{ml}}{250\text{ml}} \times 100 = 0.2\%$$

Burette:

$$\frac{0.05\text{ml}}{25\text{ml}} \times 100 = 0.2\%$$

Scales:

$$\begin{aligned} \frac{0.0005\text{g}}{40\text{g}} \times 100 &= \frac{1}{800}\% \\ \frac{0.0005\text{g}}{40.5\text{g}} \times 100 &= \frac{1}{810}\% \end{aligned}$$

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Pictures:

- Page 1: http://en.wikipedia.org/wiki/File:Galvanic_Cell.svg
- Figure 1,3: Based on Further Advanced Chemistry, figure 3.10, pg. 37
- Figure 2: Based on Further Advanced Chemistry, figure 3.11, pg. 38
- Page 5: http://www.ngfl-cymru.org.uk/vtc/ngfl/science/103_new/asc1/hazardsymbol.htm