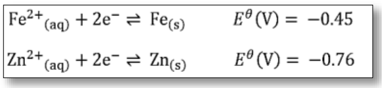
1. Research Question: **How does increasing temperature (298.15±0.50K, 303.15±0.50K, 308.15±0.50K, 313.15±0.50K, 318.15±0.50K) impact the cell potential (E_{cell}) (V), and hence equilibrium constant (K_c), in the iron-zinc galvanic cell?** The aim is to explore the nature of the relationship between temperature and E_{cell}, and temperature and K_c in a galvanic cell. The investigative approach involves constructing an iron-zinc galvanic cell from iron and zinc half cells. Cell Potential (E_{cell}) (V) is the "propulsion of electrons" in a cell measured in volts (Barrett, Navarro and Koressel) and the equilibrium constant (K_c) represents the equilibrium ratio of products to reactants. The net reaction explored is: $\mathrm{Fe^{2+}}_{(aq)} + \mathrm{Zn}_{(s)} \rightarrow \mathrm{Zn^{2+}}_{(aq)} + \mathrm{Fe}_{(s)}$.

2. Introduction: Lithium-iodide and lithium-ion batteries are galvanic cells storing chemical energy, where lithium is oxidised (Zumdahl, Zumdahl and DeCoste). Researchers are exploring replacements for lithium due to natural resources like freshwater being exhausted in extraction processes, as well as health and safety risks in usage (Grossman). An IIT Madras study illustrated the fruitful nature of utilising iron-based cells instead (Prasad). When the iron alloy steel is present under pure carbon, it can be oxidised like lithium, however since iron's reduction potential is greater than lithium, researchers are exploring the use of iron cathodes in galvanic cells (Prasad). Xu and Wang's study suggests an alternative to a lithium reducing agent is zinc; zinc-ion cells are described as having "high safety nature, low cost, and [high] environmental friendliness" (Xu and Wang) where zinc is oxidised. The iron-zinc galvanic cell has been researched to an extent. Xie, Duan and Xu have proposed a "neutral zinc–iron flow battery" that has "high energy density for stationary energy storage" and "low-cost" (Xie, Duan and Xu). These traits make this cell an ideal replacement for lithium-iodide batteries in artificial pacemakers, which needs to store and release electrical impulses over intermittently. These are electrical devices maintaining the myogenic contractions of the heart by sending out electrical impulses (Markou and Theophilidis). The internal conditions of the human body demand that pacemakers withstand a range of temperatures, to avoid malfunction. If iron-zinc batteries were to be used in pacemakers, it is important to research the cell potential to see the magnitude of the propagation of an electromotive force under different conditions. Furthermore, a galvanic cell illustrating a voltage evidences the system is not being in equilibrium. By researching K_c and temperature, researchers can make informed decisions as to the concentrations to use in cells to increase battery life (i.e., keep the reaction quotient (Q_c) far from the K_c initially so a long period of time is taken for $Q_c = K_c$ which is when no voltage is produced). It is acknowledged that this is only a small step in the wider research for the replacement of lithium-based batteries, however it is contemporarily relevant to carry out this investigation as certain literature point out research in "high temperature is still lacking and researchers need to put more efforts to promote its development" (Xu and Wang).

3. Background – 3.1 Galvanic Cells: Galvanic cells (also known as voltaic cells) generate E_{cell} from spontaneous reduction-oxidation (redox) reactions due to the movement of electrons. Half cells are made by depositing a solid metal into a solution of its own metal ions (Brown and Ford). The relationship between the solid metal and the solution of its metal ions is represented by standard electrode potentials (V). These values are calculated when the half-cell is attached to a standard hydrogen electrode (SHE), which has a half cell potential of 0V at 298K (Brown and Ford). The standard electrode potential provides a quantitative measure of the reduction potential of the metal ion in question. The more positive a standard electrode potential, the greater the reduction potential. In a galvanic cell, the half-cell with the more positive standard electrode potential will be cathode and site of reduction whilst the other half cell will serve as the anode and site of oxidation. For these individual half cells to become a galvanic cell, the 'circuit' must be completed. A salt bridge is therefore required; it is a bent piece of absorptive material soaked in a solution of salt. The salt bridge ions move to either half-cell to stop a net charge building up, it should be noted the salt should be chosen as to not interfere with the reactions at the anode and cathode. The overall cell potential of the galvanic cell is found by connecting the electrical wires to a voltmeter, or in the instance of this investigation, a Multimeter.

3.2 Iron-Zinc Galvanic Cells: The reversible reactions (right) illustrate the reduction of the metals in the half cells, because $\mathrm{Fe^{2+}}$ has a more positive electrode potential value, it will be the more readily reduced species and hence this experiment has an iron cathode and zinc anode.



$\mathrm{Fe^{2+}}_{(aq)} + 2e^- \rightleftharpoons \mathrm{Fe}_{(s)} \quad E^{\theta}(V) = -0.45$

$\mathrm{Zn^{2+}}_{(aq)} + 2e^- \rightleftharpoons \mathrm{Zn}_{(s)} \quad E^{\theta}(V) = -0.76$

Figure 1 - Standard electrode potential values at 298K (International Baccalaureate Organization)

Half-Reaction at the cathode (Reduction)	$\mathrm{Fe^{2+}}_{(aq)} + 2e^- \rightleftharpoons \mathrm{Fe}_{(s)}$
Half-Reaction at the anode (Oxidation)	$\mathrm{Zn}_{(s)} \rightleftharpoons \mathrm{Zn^{2+}}_{(aq)} + 2e^-$
Net reaction (Forward)	$\mathrm{Fe^{2+}}_{(aq)} + \mathrm{Zn}_{(s)} \rightarrow \mathrm{Zn^{2+}}_{(aq)} + \mathrm{Fe}_{(s)}$
Net reversible reaction	$\mathrm{Fe^{2+}}_{(aq)} + \mathrm{Zn}_{(s)} \rightleftharpoons \mathrm{Zn^{2+}}_{(aq)} + \mathrm{Fe}_{(s)}$
Cell diagram convention	$\mathrm{Zn}_{(s)} \vert \mathrm{Zn^{2+}}_{(aq)} \Vert \mathrm{Fe^{2+}}_{(aq)} \vert \mathrm{Fe}_{(s)}$

Table 1 – Relevant chemical equations to the investigation, and cell diagram.

The solution at the cathode and anode was 1.0M ferrous (II) sulphate and 1.0M zinc sulphate, respectively. The salt bridge, represented by the double-vertical line on the cell diagram, is 1.0M potassium nitrate solution. Galvanic cells are spontaneous and exothermic (Columbia University).

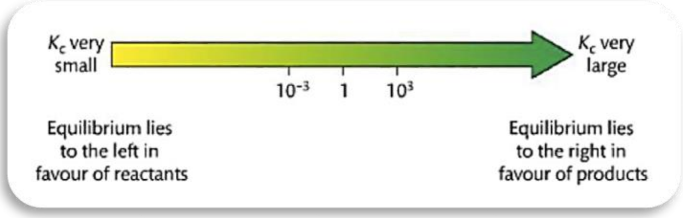


Figure 2 - Magnitude of Kc on a number line (Brown and Ford)

3.3 Reaction Quotient and Equilibrium Constant: K_c is the equilibrium constant, it represents the equilibrium ratio of products to reactants. Its value can only be affected by temperature. Its magnitude suggests whether the equilibrium lies greater on the product or reactant side (fig.2). To find the K_c, the reaction quotient (Q_c) is needed, which represents the ratio between products and reactants in a chemical reaction. Q_c has the same equation as K_c, $\frac{[C]^c[D]^d}{[A]^a[B]^b}$ (LibreTexts). For a heterogenous mixture like the iron-zinc galvanic cell only reactants or products that are not pure liquids or solids are used in the equation. Equimolar solutions were used, since these solutions were previously made by the technician, the uncertainty had already been propagated and provided for both (table 2). Though it is understood that Q_c is not a constant and the system will change concentrations by favouring the forward or reverse reaction to reach the K_c value, this experiment does assume it is constant to some extent since it is replaced every set of trials, with each set of trials taking less than 20 seconds. There is potentially a great systematic error, if the figure does not deviate from $1.00 \pm 5.9 \times 10^{-2}$, however, to adhere to green chemistry principles and lack of chemicals available, this decision was made. Furthermore, it can be categorically stated, that if at Q_c and K_c = $1.00 \pm 5.9 \times 10^{-2}$, then E_{cell} will be 0V.

$Q_c = \frac{[C]^c[D]^d}{[A]^a[B]^b} = \frac{[\text{Zn}^{2+}{}_{(\text{aq})}][\text{Fe}_{(\text{s})}]}{[\text{Zn}_{(\text{s})}][\text{Fe}^{2+}{}_{(\text{aq})}]} = \frac{[\text{Zn}^{2+}{}_{(\text{aq})}]\cancel{[\text{Fe}_{(\text{s})}]}}{\cancel{[\text{Zn}_{(\text{s})}]}[\text{Fe}^{2+}{}_{(\text{aq})}]} = \frac{[\text{Zn}^{2+}{}_{(\text{aq})}]}{[\text{Fe}^{2+}{}_{(\text{aq})}]} = \frac{[1.00 \pm 4.78 \times 10^{-2}\ \text{moldm}^{-3}]}{[1.00 \pm 1.01 \times 10^{-2}\ \text{moldm}^{-3}]}$		
Maximum value: $\frac{[1.0478]}{[0.9899]} = 1.0590$	Measured value: $\frac{[1.00]}{[1.00]} = 1.00$	Minimum value: $\frac{[0.9522]}{[1.0101]} = 0.9427$
Absolute uncertainty: $1.059 - 1.00 = \pm 5.90 \times 10^{-2}$		
$Q_c = 1.00 \pm 5.90 \times 10^{-2}$		

Table 2 - Absolute uncertainty of the reaction quotient calculation, reaction quotient formula (LibreTexts)

3.4 Derivation of the Equilibrium Constant Equation: Gibbs free energy change conveys the spontaneity of a reaction, in an exothermic reaction such as this one. From Gibbs free energy change, the equilibrium constant K_c can be found. To do this, the Gibbs free energy change must be in its standard state, hence the Nernst equation and the equation for Gibbs free energy and cell potential can be combined. This combined equation can be equated with the standard Gibbs free energy and equilibrium equation and rearranged to illustrate an equation for the equilibrium constant K_c.

Name	Equation	Terms of the Equation
*Equation 1 - Change in Gibbs free energy **	$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$	$\Delta G^{\ominus} = change\ in\ Gibbs\ free\ energy\ at\ standard\ conditions$ $\Delta H^{\ominus} = change\ in\ Enthalpy\ at\ standard\ conditions\ *$ $\Delta S^{\ominus} = change\ in\ Entropy\ at\ standard\ conditions\ *$ $T = Temperature\ in\ Kelvin$
*Equation 2 - Nernst Equation**	$E = E^{\ominus} - \left(\frac{RT}{nF}\right)\ln Q$	$E = Cell\ potential$ $E^{\ominus} = Cell\ potential\ \ at\ standard\ conditions\ *$ $n = Moles\ of\ electrons\ involved\ in\ 1\ mole\ of\ reaction$ (Brown and Ford) $F = Faraday's\ constant = 9.65 \times 10^4(\text{Cmol}^{-1})\ **$ $R = Gas\ constant = 8.31(\text{JK}^{-1}\text{mol}^{-1})\ **$ $ln\ Q = the\ natural\ \log of\ Q_c$
*Equation 3 – Standard Gibbs free energy and standard cell potential**	$\Delta G^{\ominus} = -nFE^{\ominus}$	
*Equation 4 - Standard Gibbs free energy and equilibrium constant**	$\Delta G^{\ominus} = -\text{RT}\ln K$	$ln\ K = the\ natural\ \log of\ K_c$

*Table 3 - equations to be combined in table 3, *Taken from the Chemistry Data Booklet (2014) (International Baccalaureate Organization)*

Step Number	Step Description	Equation
1	Make ‘$E^{\ominus}$’ the subject in equation 2	$E^{\ominus} = E + \left(\frac{RT}{nF}\right)\ln Q$
2	Substitute equation 2 into equation 3	$\Delta G^{\ominus} = -nF\left[E + \left(\frac{RT}{nF}\right)\ln Q\right]$
3	Expand the brackets	$\Delta G^{\ominus} = -nFE - \mathrm{RT}\ln Q$
4	Equate the simplified equation with equation 4	$-\mathrm{RT}\ln K = -nFE - \mathrm{RT}\ln Q$
5	Divide both sides by ‘$-RT$’	$\ln K = \frac{nFE}{\mathrm{RT}} + \ln Q$
6	Take the natural log of both sides	$K_c = \ln\left(\frac{nFE}{\mathrm{RT}}\right) + Q_c$
$K_c = \ln\left(\frac{nFE}{\mathrm{RT}}\right) + Q_c$ *Equation 5 - Equilibrium constant equation*		

Table 4 – derivation of equilibrium constant equation

3.4 Le Chatelier’s Principle, Temperature and Hypothesis: Le Chatelier’s Principle states at equilibrium, a system will respond to changes by minimizing it (Brown and Ford). In an exothermic reaction, a temperature increase will theoretically shift the position of equilibrium to the left by favouring the backward reaction, the K_c reflects this by a decrease in value. This shift to the left happens because an increase in temperature (also known as absolute temperature) increases the energy within a system, and since the backward reaction is endothermic and can minimise this increase in energy, it will be favoured. Since the forward reaction is the movement of electrons from the anode to cathode, an electromotive force is created, shown by a positive E_{cell} (V) value on a voltmeter. The backward reaction is the drawing of electrons from the iron electrode to the zinc electrode, in which the electromotive force acts in the opposite direction. If the position of equilibrium moves to the left, the relative extent at which electrons is being drawn into the anode increases, hence the value of E_{cell} (V) decreases. It can therefore be hypothesised, increasing temperature (298.15±0.50K, 303.15±0.50K, 308.15±0.50K, 313.15±0.50K, 318.15±0.50K) will to some extent impact E_{cell} and K_c, in the iron-zinc galvanic cell, by a decrease in both.

4. Variables:

Independent	Possible Impact on investigation	Method of Changing and Measuring the Variable
Temperature (K) of the galvanic cells	Increasing temperature causes decrease in both E_{cell} (V) and K_c (theory in 3.4).	The cell will be placed in an electric water bath, which maintains the temperature. Progressively adding boiling water from an kettle increases the temperature from 298.15±0.50K to 318.15±0.50K in increments of 5.00±0.50K. Measured using a mercury thermometer (±0.50K).
Dependent	**Link to Research Question**	**Method of Measurement**
Voltage (V)	E_{cell} is measured in volts and can be further manipulated mathematically to provide K_c.	DM83D Digital Multimeter (±0.001V) attached to the cell in parallel, values recorded with pen and paper on a results table.
Controlled	**Possible Impact on investigation**	**Method of Control**
Galvanic cell structure (use of wiring)	The electrical wires will create resistance to some extent (opposition to the flow of electrons), which may affect E_{cell} values, for fairness this resistance must be uniform across all trials. Lack of control potentially leads to systematic error.	Using the same electrical wires and alligator clips in the entirety of the experiment, they will be marked using a whiteboard pen so that they are used in every trial.
Volume and of solutions	If the volume of the metal solution in each half cell were to be less than 100cm^3 then the electrode would not be submerged completely during each trial. The surface	Using measuring precise instruments (measuring cylinder 50cm^3±5 cm^3) to

Volume and of solutions (continued)	area available for reduction and oxidation purposes would decrease and hence the outputted value would not be accurate, thereby potentially contributing to systematic error.	make and use solutions of intended concentration and volume.
Concentration of solutions	If the concentration of metal solutions were to be values other than 1.00M then Q_c could not be assumed to be $1.00 \pm 5.90 \times 10^{-2}$, increasing systematic error in calculation thereby increasing inaccuracy of analysis and conclusion.	Discussing the precision involved of previously made solutions with the lab technician. Marking the concentration of solutions with a whiteboard pen
Electrodes	If the same electrodes were kept the same, the E_{cell} readings will not illustrate the true values, as the surface area of electrodes for redox reactions change over time due to material being added and lost. Oxides that have formed on the surface prior to being in the cell can also impact the surface area in the same way.	Replacing electrodes at every increment of temperature and cleaning the electrodes with sandpaper to remove oxides.
Time for each set of trials	If the time per trial is too long, then Q_c will change significantly. In the spirit of green chemistry, and due to constraints in chemical availability, this investigation expects to reuse solutions and by doing so assume Q_c=1	Using a timer, making sure that data for 5 trials per temperature take less than 20 seconds to obtain. After 20 seconds the salt bridge will be removed to stop the redox reactions.
Uncontrolled	**Possible Impact on Investigation**	**Method of Reducing Impact**
Changes in reaction quotient (Q_c)	Because Q_c will change concentrations to become the equilibrium constant value, it is entirely possible that this change can happen quickly leading to major systematic errors in calculation and thus inaccurate conclusions.	a timer is used, making sure that data for 5 trials per temperature take less than 20 seconds to obtain. After 20 seconds the salt bridge will be removed to stop the redox reactions and the solution is replaced.

5. Preliminary Testing and Justification: The method used in preliminary testing was obtained from *C2507 Intensive General Chemistry – E10 Electrochemistry and Thermodynamics* (Columbia University) and adapted to suit the needs of the investigation. Adaptations include changing the range and increments of temperature. The original method stipulated a low temperature range (with the lowest value being 288.15K±0.50K), this was not feasible, because a source of low temperature was not available to cool water. It was modified to 298.15±0.50K to 318.15±0.50K, this modification was also made in context of the investigative purpose, as this is the temperature at which pacemakers operate and research in "high temperature is still lacking" (refer to 1. Introduction) (Xu and Wang). In preliminary testing, the candidate sustained injury using the hotplate. A water bath was used instead, a greater sense of caution was adopted during the data collection by wearing a lab coat. This improvement reduces systematic errors, as the available hotplate does not have a scale. To improve the reliability of the data produced, an effort was made to include 5 repeat readings at each increment of independent variable. Furthermore, the decision was made to record qualitative data as well. This would help substantiate or disprove claims made in the analysis or conclusion. Lastly, 100cm^3 was used for metal solutions in each half cell rather than the 400cm^3 stipulated in the methodology to adhere to atom economy in green chemistry (Cann) whilst making sure at least three-quarters of the electrodes were submerged. The advantage of adapting a method is that it leads to less systematic error, as it has perceivably been reproduced and improved by professionals prior to publication due to the credibility of the source (Columbia University).

6. Apparatus: **_Since this is a procedural step, rather than the measurement of time, there is no uncertainty._

Chemicals and expendable goods	Glassware / Containers	Lab Equipment	
500cm^3 of iron (II) sulfate $1.00 \pm 1.01 \times 10^{-2}$ moldm^{-3} 500cm^3 of zinc sulfate $1.00 \pm 4.78 \times 10^{-2}$ moldm^{-3} 300cm^3 of potassium nitrate $1.00 \pm 1.19 \times 10^{-2}$ moldm^{-3} 5 x iron nails 1 x zinc foil sheet Source of tap water	2 x Mercury Thermometer (±0.5K) 3 x 50±5cm^3 Measuring cylinder 3 x 100cm^3 Beakers 1 x 2dm^3 Water jug	1 x DM83D digital Multimeter (±0.001V) 1 x Scissors 1 x Ruler (±0.05cm) 1 x Sandpaper sheet 4 x Filter paper 2 x Crocodile clips 2 x Electrical wires 1 x Safety goggles 1 x Lab coat	1 x Electric water bath 1 x Kettle 1 x Phone (for timer application)** 1 x Glass food container 1 x Tongs 1 x Chemical waste bin 1 x Rubber Gloves 1 x Mask 1 x Pen and Paper

7. Method: Adapted from *C2507 Intensive General Chemistry – E10 Electrochemistry and Thermodynamics* (Columbia University). The direct objective of the method is to find E_{cell} (V). Rubber gloves, safety goggles and Lab coat should be worn throughout. *N.B. Mercury based thermometer readings, measured in degrees Celsius (°C), were converted to Kelvin (K) by the addition of 273.15K (Lumenlearning).*

Part I. Set Up

1. Using two 50±5cm^3 measuring cylinders, measure out 50cm^3 of zinc sulphate and 50cm^3 of iron sulphate.
2. Add each solution to a separate 100cm^3 beaker.
3. Fill the 2dm^3 measuring jug with water from the tap and pour into the electric water bath .
4. Fill the kettle with water until
5. Add the boiled water to the kettle until 298.15K is reached
6. Place an inverted food container on the floor of the water bath
7. Place both metal sulphate solutions side by side on top of the container
8. Place a thermometer in each solution

N.B at this point, half of the beakers should be submerged in the water

9. Start a stopwatch for 5 minutes for complete thermal equilibrium to be reached at 298.15±0.50K, check the needed value has been reached
10. Behind the water bath, place a DM83D Multimeter
11. Attach two electrical wires to the D.C. ports of the Multimeter
12. Attach crocodile clips to the ends of both electrical wires
13. Using the 50±5cm^3 measuring cylinder, measure out 100cm^3 of potassium nitrate.
14. Add the potassium nitrate to a 100cm^3 beaker and place it to one side

Part II. Data Collection

1. Attach an iron nail to one of the crocodile clips
2. Measure 3.00±0.05cm by 5.00±0.05cm using the Ruler (±0.05cm), on the zinc foil sheet
3. Use scissors to cut out the zinc foil according to the measurements
4. Attach this piece of zinc foil to the other crocodile clip
5. Clean both pieces of metal (electrodes) with sandpaper
6. Place the zinc electrode into the solution of zinc sulphate.
7. Place the iron electrode (iron nail) into the solution of iron (II) sulphate
   N.B. at least three quarters of the electrodes should be submerged
8. Measure 10.00±0.05cm by 3.00±0.05cm of filter paper using a ruler
9. Cut the strip of filter paper out using scissors.
10. Soak the strip of filter paper in the potassium nitrate using tongs
11. Start a timer for 20 seconds on your phone
12. Place the strip (salt bridge) in between the two half cells in the water baths, making sure both ends are touching the solutions.
13. Turn on the DM83D Multimeter and turn the dial to the voltmeter setting
14. Record the value of the voltage on a data table on a piece of paper
15. Turn Multimeter dial to original dial and turn off instrument, turn instrument on again, record five values over five repeats
16. Remove the salt bridge with tongs and dispose in the chemical waste bin
17. Dispose contents of the half cells in the chemical waste bin and replace with the respective solution, using 50±5cm^3 for measurement
18. Fill the kettle with tap water and boil
19. Add boiling water to the water bath until the next temperature is reached, allow 5 minutes using a timer for thermal equilibrium to be reached
20. Repeat steps 1 to 19 for the temperatures: 303.15±0.50K, 308.15±0.50K, 313.15±0.50K, 318.15±0.50K

Part III. Post Data Collection

The contents of the half cells should be disposed of in the chemical waste bin located in the fume cupboard immediately. Turn off power source to water bath and empty into the sink. Unused chemicals or uncontaminated solutions should be returned to storage.

8. Safety, Ethical and Environmental Concerns:

Concern	Note
Iron (II) sulfate solution	Iron (II) and zinc sulphate Irritant to the eyes and skin, harmful if swallowed (CLEAPPS), hence a mask, gloves and safety goggles will be worn. Potassium nitrate causes skin and eye irritation (CLEAPPS), irritates the “nasal cavity and thorax” (New Jersey Department of Health and Senior Services). Experiment will also be conducted in the fume cupboard.
Zinc sulfate solution	
Potassium nitrate solution	
High temperature water-bath	Safety goggles worn to protect eyes from steam and boiling water, gloves worn and contact actively minimised with the water due to potential of scalding.
Ethical and Environmental awareness	Solutions used in the experiment are only used with necessity as to not waste chemicals and to adhere to green chemistry principles like atom economy (Cann). After a day of experimentation, on the subsequent day the chemicals can stay the same but must be wrapped in clingfilm and stored in a secure area. Once the experimentation is completed, chemicals can be thrown away in the chemical waste bin where they will be disposed of properly. Ethically speaking, animals or other students were not in immediate danger from apparatus or chemicals used, as the experimental setup was deconstructed immediately after use and chemicals were stored in a secure area.

9. Results – 9.1 Raw data:

Temperature (±0.5K)	**E_{cell} (±0.001V)**					**Average E_{cell} (±0.001V)**
	Trial Number					
	1	2	3	4	5	
298.15	~~0.420***~~	0.429	0.429	0.429	0.429	0.429
303.15	0.409	0.407	0.406	0.406	0.406	0.407
308.15	0.398	0.398	0.398	0.398	0.397	0.398
313.15	0.393	0.393	0.393	0.393	0.393	0.393
318.15	0.388	0.388	0.388	0.388	0.389	0.389

Table 5 - Cell potential (V)

***Outlying piece of data excluded from the average E_{cell}.

9.2 Qualitative data: After each set of trials, when the iron cathode and zinc cathode were replaced, the smoothness that both electrodes previously had, was very slightly diminished, a slightly rougher surface than before was felt. This occurred at all temperatures apart from 318.15±0.50K. Differences in roughness between sets of data was not found, but perhaps this could have been seen to a greater extent had the cell ran for a longer period. The smoothness of the electrodes had diminished because of material build-up on the sides of the electrodes. On the iron electrode, electrons are being gained, hence Fe^{2+} ions in the ferrous (II) sulphate solution move toward the electron dense cathode and attaches. The zinc electrode loses mass as electrons are being lost, hence Zn^{2+} ions will be formed and move to nitrate ions from the potassium nitrate solution (salt bridge).

9.3 Calculation Examples:

Average E_{cell} (V)	$\frac{0.429+0.429+0.429+0.429}{4} = 0.429$
Average E_{cell} (V) uncertainty	$\frac{Maximum - Minimum}{2} = \frac{0.429-0.429}{2} = \pm 0$
Absolute Uncertainty for Average E_{cell} (V)	$Half\ Range\ Uncertainty + Measurement\ uncertainty$ $= 0 + 1.00 \times 10^{-3} = \pm 1.00 \times 10^{-3}$

Table 6 - Average Ecell (V) values and Uncertainty Using Half-Range Method

Equation 6 **Average K_c Combined Equation**	$K_c = \ln\left(\frac{nFE}{RT}\right) + Q_c$
Substitute the value for Faraday's constant F = 96500 C mol^{-1} (International Baccalaureate Organization)	$K_c = \ln\left(\frac{n \times 96500 \times E}{RT}\right) + Q_c$
Substitute the value for the gas constant R = 8.314 $JK^{-1}mol$ (International Baccalaureate Organization)	$K_c = \ln\left(\frac{n \times 96500 \times E}{8.314 \times T}\right) + Q_c$
Substitute the number of moles of electrons in transfer, 2 electrons are transferred in this reaction, therefore, n = 2.00	$K_c = \ln\left(\frac{2.00 \times 96500 \times E}{8.314 \times T}\right) + Q_c$
Substitute the value for the reaction quotient, $Q_c = 1$ (Refer to section 3.3)	$K_c = \ln\left(\frac{2.00 \times 96500 \times E}{8.314 \times T}\right) + 1.00$
Substitute the value 298.15K for T	$K_c = \ln\left(\frac{2.00 \times 96500 \times E}{8.314 \times 298.15}\right) + 1.00$
Substitute the value for average E_{cell} 0.427V	$K_c = \ln\left(\frac{2.00 \times 96500 \times 0.429}{8.314 \times 298.15}\right) + 1.00$
Natural Log of terms in brackets	$= \ln(32.3) + 1.00$ $= 3.48 + 1.00$
Sum	$= 4.48$
Equilibrium Constant, K_c	$K_c = 4.48$

Table 7 - Equilibrium constant combined equation example

	Value	Maximum Value	Minimum Value
Equilibrium Constant, K_c	4.48	$\ln\left(\frac{2.00 \times 96500 \times 0.433}{8.314 \times 297.65}\right) + (1.059) = 4.54$	$\ln\left(\frac{2.00 \times 96500 \times 0.422}{8.314 \times 298.65}\right) + (0.9427) = 4.43$

Table 8 – Example maximum and minimum value calculation for equilibrium constant

	Value	Absolute Uncertainty (Maximum – Recorded Value) (as suggested by Brown and Ford)	Percentage uncertainty (%)
Equilibrium Constant, K_c	4.48	$4.54 - 4.48 = 6.30 \times 10^{-2}$	$\left(\frac{6.30 \times 10^{-2}}{4.48}\right) \times \frac{100}{1} = 1.41 \times 10^{0}\ \%$

Table 9 - Absolute Uncertainty and Percentage Uncertainty Calculation

9.4 Processed data:

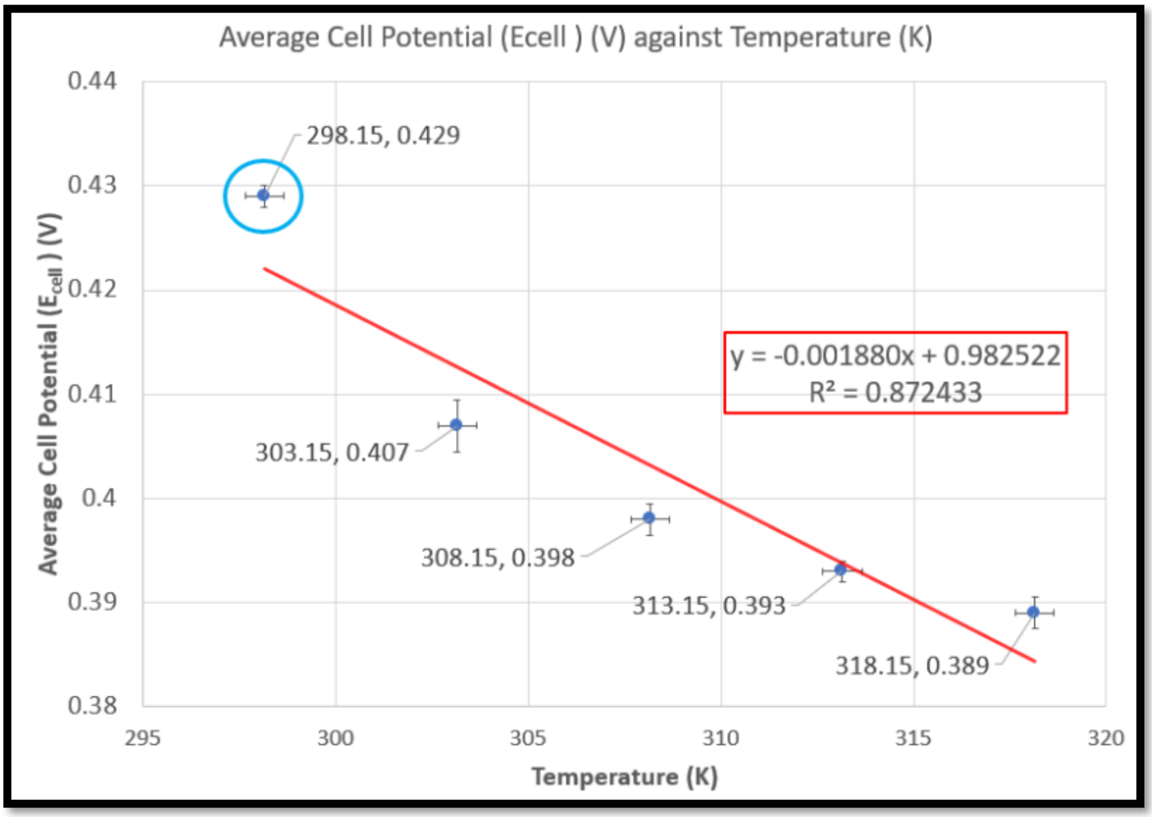
Temperature (±0.5K)	Average E_{cell} (±0.001V)	Average E_{cell} uncertainty (V)	Average E_{cell} Total uncertainty (V)	Percentage uncertainty (%)
298.15	0.429	-	$\pm 1.00 \times 10^{-3}$	2.31×10^{-1}
303.15	0.407	$\pm 1.50 \times 10^{-3}$	$\pm 2.50 \times 10^{-3}$	6.14×10^{-1}
308.15	0.398	$\pm 5.00 \times 10^{-4}$	$\pm 1.50 \times 10^{-3}$	3.77×10^{-1}
313.15	0.393	-	$\pm 1.00 \times 10^{-3}$	2.54×10^{-1}
318.15	0.389	$\pm 5.00 \times 10^{-4}$	$\pm 1.50 \times 10^{-3}$	3.85×10^{-1}

Table 10 – Processed data for E_{cell}

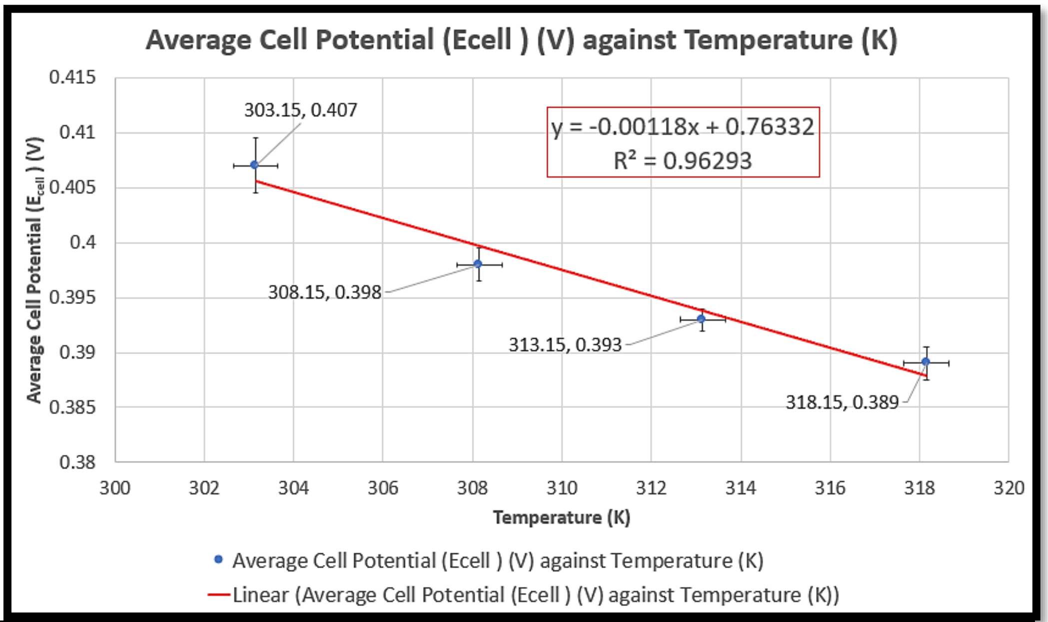
Temperature (±0.5K)	Average K_c	Absolute uncertainty (Maximum – Recorded Value)	Percentage uncertainty (%)
298.15	4.51×10^{0}	6.30×10^{-2}	1.41×10^{0}
303.15	4.43×10^{0}	6.68×10^{-2}	2.01×10^{0}
308.15	4.40×10^{0}	6.43×10^{-2}	1.46×10^{0}
313.15	4.37×10^{0}	6.31×10^{-2}	1.44×10^{0}
318.15	4.35×10^{0}	6.44×10^{-2}	1.48×10^{0}

Table 11 - Processed data for Kc

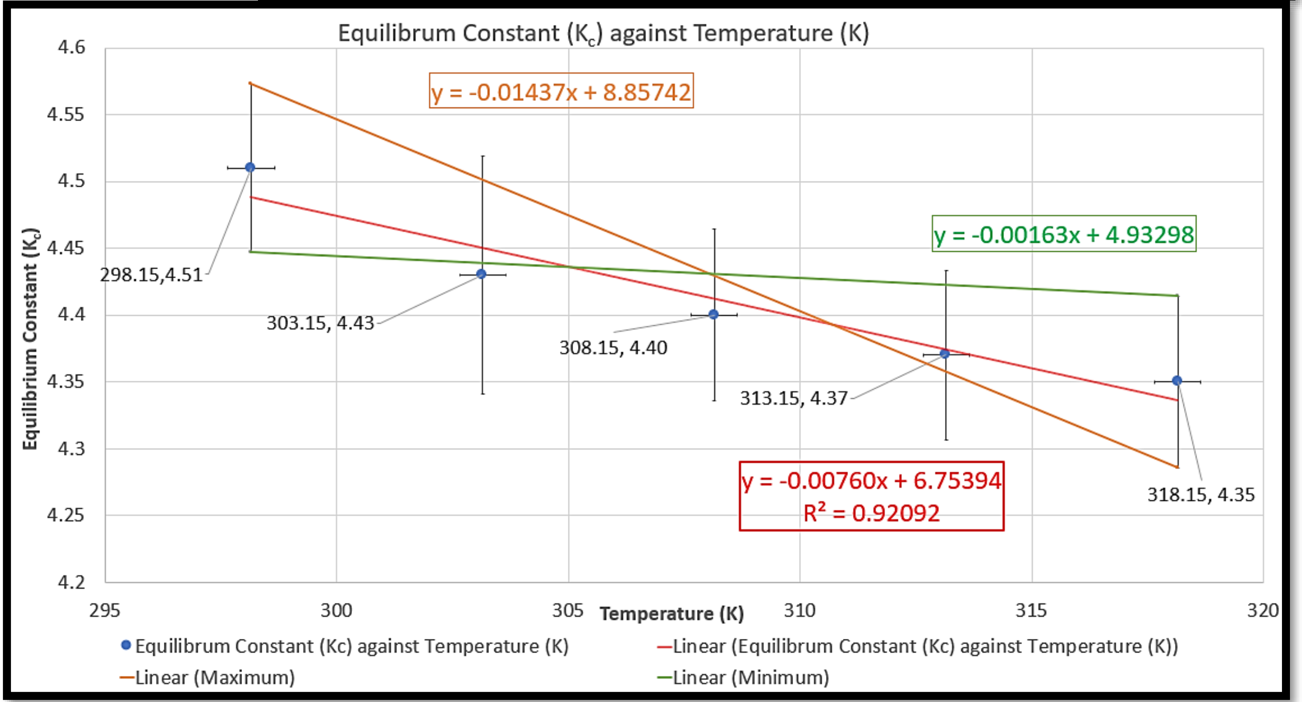
10. Data Presentation:



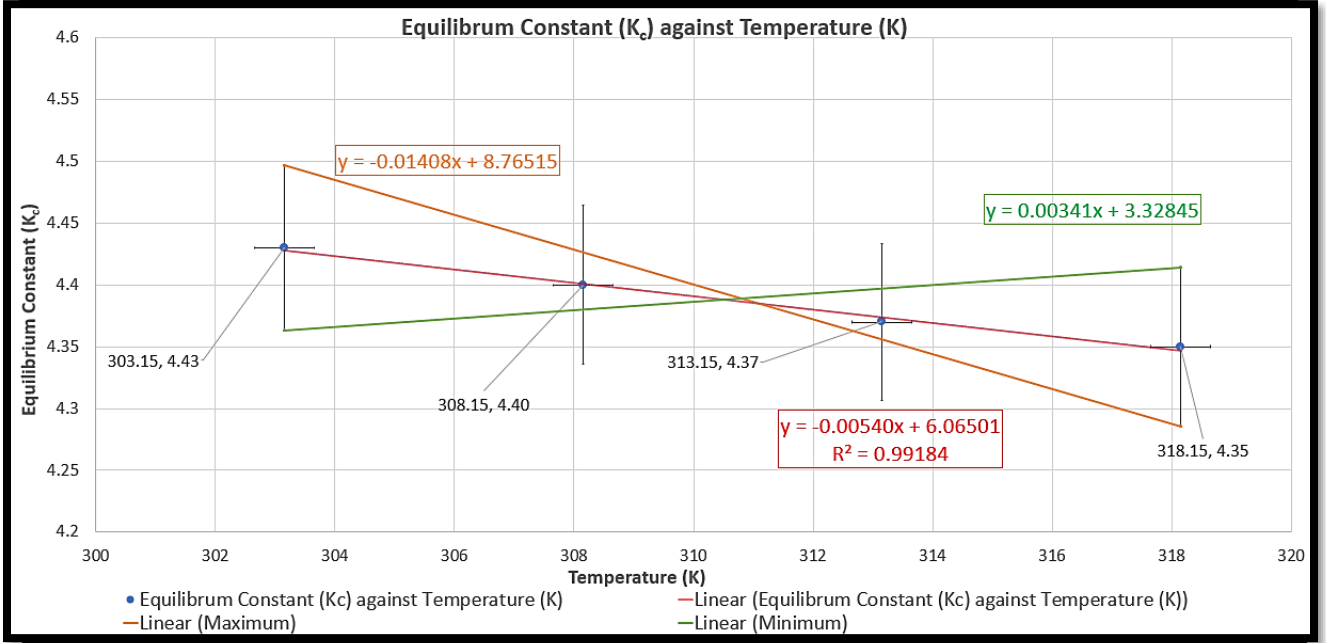
Graph 1 – line of best fit of Average K_c results in table 10, with horizontal bars representing temperature uncertainty, vertical bars represent absolute uncertainty of E_{cell}



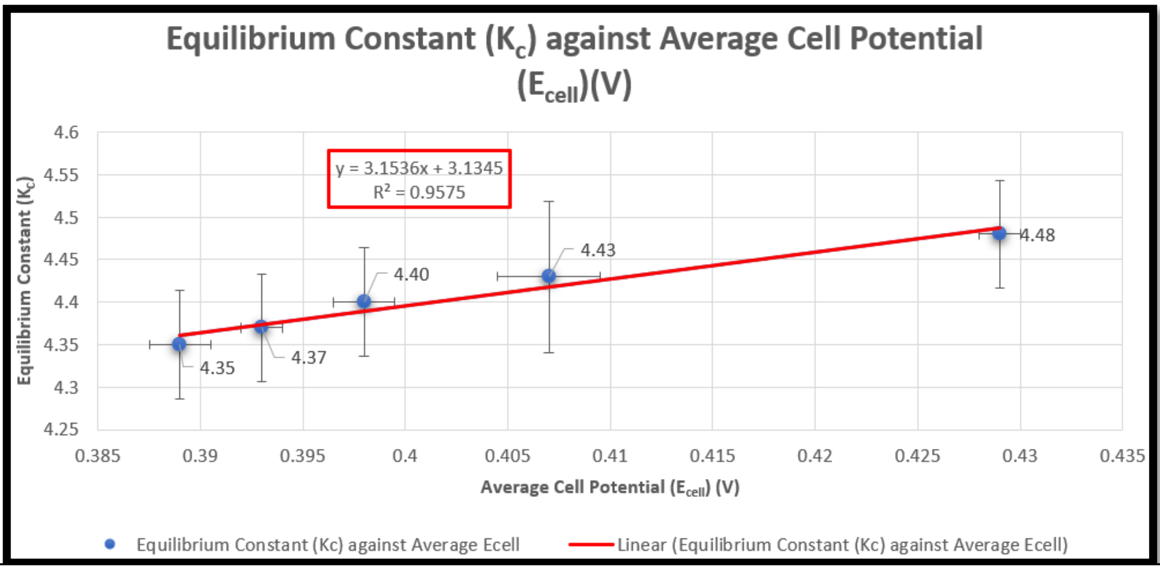
Graph 2 – line of best fit of Average K_c results in table 10 without the 298.15±0.50K data point, with horizontal bars representing temperature uncertainty, vertical bars represent absolute uncertainty of E_{cell}



Graph 3 – Line of best fit of results of Table 11, with vertical bars representing individual absolute uncertainty of K_c, horizontal bars representing temperature uncertainty.



Graph 4 – Line of best fit of results of Table 11 without the 298.15±0.50K data point, with vertical bars representing individual absolute uncertainty of K_c, horizontal bars representing temperature uncertainty.



Graph 5 - Line graph and line of best fit of results of Table 10 and 11, with vertical bars representing individual absolute uncertainty of K_c, horizontal bars representing absolute uncertainty of E_{cell}

11. Analysis - 11.1 Trends and Supporting Theory: Graph 1 illustrates a negative correlation between increasing temperature and Average E_{cell}. A polynomial trend or linear trend could be applied to the data. When referring to literature, studies on E_{cell} values conducted in similar temperature ranges (293.15 – 313.15K) (Jakubowska) and conducted in different ranges (400 – 600K) (Kuterbekov, Balapanov and Ishembetov) suggest that a negative linear correlation exists between average E_{cell} and temperature. A linear line of best fit is thereby applied to the data on in table 10 represented by the red line on graph 1, the equation is $y = -0.001880x + 0.982522$. The negative coefficient of the x suggests, with increasing 'x' value, the 'y' value will become numerically smaller. Pearson's Correlation Coefficient can be used to find out the strength of the relationship represented by this equation. The closer an R^2 value is to 1, the stronger the correlation. The R^2 value of trendline is 0.872, suggesting that when represented as a linear relationship, the correlation is strong. Strong correlations are illustrated by R^2 values greater than 0.750 (Francis). Alternatively, a curve of best fit could be applied. This curve would be represented by the polynomial (quadratic) equation $y = 0.0001x^2 - 0.0723x + 11.829$ and may represent correlation better as the R^2 value is 0.986, representing a stronger correlation between the independent and dependent variable. A curve suggests that the rate of change, or the extent at which temperature affects average E_{cell} is not constant and decreases with more cumulative temperature increase – this is seen mathematically, as the equation contains a variable with an exponent, x^2. However, the problem with taking a quadratic equation as a curve of best fit, is that it suggests at if this relationship was continued, at some point the cell potential will increase after having decreased over increased temperature – since no literature was found to support this idea, this curve is not a good representation of the relationship, therefore the linear equation will be used for further analysis. Graph 3 shows a negative correlation between increasing temperature and K_c because its equation has a negative coefficient on x. The trendline suggests that a negative linear correlation exists between K_c and temperature has an R^2 value of 0.9909 the correlation is strong. For the same reasons as graph 1, the linear line of best fit will be taken for further analysis instead of a curve of best fit.

Both negative correlation trends in graphs 1 and 3 are supported by scientific theory. A galvanic cell is always exothermic and spontaneous, therefore when applying an increase in temperature, according to Le Chatelier's Principle, equilibrium shifts to the left favouring endothermic reaction to a greater extent. By doing so the changes in heat to the system are minimized however the reactants are favoured, this is reflected in the average E_{cell} and K_c. The voltage created in the endothermic reaction moves opposite to the exothermic reaction, the flow of electrons is from the iron to the zinc electrode instead. This opposes the flow of electrons from the anode to cathode, the voltage propagating in both directions cancel out, thereby decreasing the overall value of the E_{cell}. For the K_c, since equilibrium shifts to the left, a greater proportion of reactants will be present to products. In the sense that the concentration of reacts will become larger and the concentration of products will become smaller at equilibrium. Thinking about this in the context of the equation, $K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$ (LibreTexts), the denominator becomes greater and the numerator becomes lesser. Therefore, the value of K_c decreases with an increase in temperature.

Seeing that both graph 1 and graph 3 display similar trends with temperature, it is important to look at the extent at which their data can be correlated. Graph 5 illustrates a linear and positive correlation between K_c and average E_{cell} in the equation $y = 3.1536x + 3.1345$. The R^2 is 0.958, a very strong display of correlation, therefore both relationships with temperature should be negative. Furthermore, K_c in the context of this investigation is calculated with the input of E_{cell} values, therefore the fact that the relationships are similar is expected.

11.2 Anomalous Results and Error Bars: Anomalous data can be detrimental to the interpretation of the results produced; therefore, it becomes important to identify its presence. In graph 1, looking at the data point at 298.15±0.50K compared to the rest of the data, it is much higher than expected due to its deviation from the trendline. Though it could be thought that perhaps there may be a changing gradient because of an exponential or polynomial relationship causing a sharp decrease in subsequent values from 298.15±0.50K, this is not the case as the established relationship is linear. Since only 5 data points exist on the graph, the linear correlation is heavily affected by even one data point. The trendline may be adhering to the 298.15±0.50K data point, shown by the lack of intersection with the error bars of 4 out of 5 points. The trendline can be seen without the 298.15±0.50K data point in graph 2. This trendline is a better representation of the relationship having removed the anomalous result, as at least one error bar on each data point intersects. Removing the anomalous piece of data has not yielded any significant change to the nature of the relationship between the variables, there is still a strongly negative trendline (equation, $y = -0.00118x + 0.76332$). The R^2 value is now 0.963, a 10.4% increase from before, suggesting the trendline has become a better representation of the correlation between average E_{cell} and temperature. Since the anomalous piece of data was the average of the 4 trials (1 anomalous trial), the random error was actively decreased through repeats, this anomalous result is likely the product of systematic error for this set of trials. Reflecting on the data collection, whilst measuring the temperature of the solutions, there was parallax error to some degree. In the sense that the value of the temperature was not taken at eye level for this set of data, it may have been thought that the value was at 298.15K but may have been lower due to thermal equilibrium not being reached. A temperature lower than 298.15K was most likely achieved but not noticed, thereby causing an elevated E_{cell} value.

For graph 3, representing K_c against temperature, the error bars are very large vertically and are small horizontally, there is significant overlap between error bars suggesting that the difference between values may have occurred by chance. Due to this large degree of overlap between multiple points, maximum and minimum trendlines are shown on graph 3 to illustrate the potential effect of this error on the entire relationship between temperature and K_c. The maximum and minimum trendline equations are $y = -0.01437x + 8.85742$ and $y = -0.00163x + 4.93298$ respectively. Although the coefficient of x differs in the two equations by close to a factor of 10, the negative coefficient still stands, suggesting that regardless of the extent of error in this data, the negative correlation between temperature and K_c exists. Having previously looked at the relationship between K_c and average E_{cell}, the 298.15±0.50K was removed. In graph 3, representing K_c against temperature, although the 298.15±0.50K data point does not seem anomalous regarding the overall trend, since it was calculated using the previous anomalous result of 298.15±0.50K in graph 1, the effect of its removal is seen on graph 4. Graph 4's R^2 value of 0.99 suggests that the new regression line (equation, $y = -0.00540x + 6.06501$) almost fits the data perfectly, the relationship is less ambiguous as every line intersects the regression line. However, the removal of the anomalous result has uncovered that the overall trend could be reversed, and that increasing temperature may increase K_c. The minimum trendline (green) has an equation of $y = 0.00341x + 3.32845$, since the coefficient of x is positive, the correlation is positive as well.

11.5 Uncertainty in Trendlines**:** Analysis has shown that increasing temperature decreases both average E_{cell} and K_c; both have negative linear correlations with the equations $y = -0.00118x + 0.76332$ and $y = -0.00540x + 6.06501$ respectively, having removed the anomalous data point at 298.15±0.50K. The error bars for temperature and average E_{cell} are small, therefore using maximum and minimum lines are not feasible to obtain error. Instead, the line estimate function on MS Excel can be employed which uses the uncertainty of the slope itself. The coefficient of x in the average E_{cell} against temperature is $-0.00118 \pm 1.64 \times 10^{-4}$. For K_c against temperature, since the error bars are larger, maximum and minimum lines (graph 4) can be used to find the uncertainty in gradient. The half range method (table 6) can be used:

$$\frac{(\text{Maximum Gradient} - \text{Minimum Gradient})}{2} = \frac{(-0.01408 - 0.00341)}{2} = \pm 0.008745 = \pm 8.75 \times 10^{-3}$$

The coefficient of x in K_c against temperature is $-0.00540 \pm 8.75 \times 10^{-3}$. Percentage uncertainty is calculated using the formula below for the gradients of average E_{cell} and temperature, and K_c and temperature, respectively:

$$\frac{\text{Uncertainty}}{\text{Measured Value}} \times 100 = \frac{1.64 \times 10^{-4}}{-0.001182} \times 100 = \pm 13.9\%$$

$$\frac{\text{Uncertainty}}{\text{Measured Value}} \times 100 = \frac{8.75 \times 10^{-3}}{-0.00540} \times 100 = \pm 162\%$$

11.6 Analysis summary: Having removed the anomalous data point at 298.15±0.50K, it has been found that increasing temperature decreases both average E$_{cell}$ and K_c; both linear correlations are negative with the equations $y = -0.00118 \pm 13.9\%x + 0.76332$ and $y = -0.00540 \pm 162\%x + 6.06501$ respectively. On the latter equation, for K_c against temperature, the percentage error was extremely high (over 100%), suggesting that the true value's deviation from the measured value is not unexpected. As a result, the minimum line propagated suggested that the relationship between K_c against temperature could be positive. Though percentage error shows that this could be the case, there is no scientific theory to corroborate this trend (refer to section 11.1) and in fact it contradicts scientific theory. Average E$_{cell}$ against temperature saw a 13.9% uncertainty in its gradient. Regardless of the extent of uncertainty, the negative relationship seemingly remains the same in any situation, it is only magnitude of the gradient that changes. From the percentage uncertainties in table 11 (for K_c), although there is no piece of data which had greater than 2%, the analysis shows that the cumulative effect of multiple small uncertainties can impact the overall gradient significantly if the difference between values on the y-axis are small.

12. Evaluation – 12.1 Comparison with Literature: The average E$_{cell}$ value at 298.15±0.50K ($0.429 \pm 1.00 \times 10^{-3}$) can be compared with E$_{cell}$ calculated from standard electrode potentials. This is because the reading was at 298.15±0.50K and cells involved had 1 moldm^{-3} solutions. Pressure is not taken into consideration as no gaseous substances are involved.

<table>
<tr><th colspan="6">Percentage Error</th></tr>
<tr><td colspan="3">The standard E$_{cell}$ value is calculated from:
$E^{\ominus}$ cell $= E^{\ominus}$ reduction $- E^{\ominus}$oxidation
(Brown and Ford)</td><td colspan="3">$E^{\ominus}$ cell = (-0.45) − (-0.76)
$E^{\ominus}$ cell = 0.31V</td></tr>
<tr><td colspan="3">Percentage Error formula</td><td colspan="3">$\frac{(\text{Literature Value} - \text{Experimental Value})}{\text{Experimental Value}} \times 100$</td></tr>
<tr><td colspan="2">Maximum
$\frac{(0.31 - 0.430)}{0.31} \times 100 = 38.7\%$</td><td colspan="2">Measured
$\frac{(0.31 - 0.429)}{0.31} \times 100 = 38.4\%$</td><td colspan="2">Minimum
$\frac{(0.31 - 0.428)}{0.31} \times 100 = 38.1\%$</td></tr>
<tr><td colspan="6">Absolute Uncertainty: Maximum – Measured = 38.7 – 38.4 = 0.3</td></tr>
<tr><td colspan="6">38.4±0.3%</td></tr>
</table>

Table 12 – calculation of percentage error between experimental and literature value

Though it could be argued that this comparison is invalidated due to the 298.15±0.50K experimental value being disregarded in analysis as an anomaly, the percentage error can still be discussed having understood that additional systematic errors were present for this data point alone (refer to 11.2). A large percentage error may exist because Q$_c$ may have not been exactly 1 (hence its previously stated uncertainty of $\pm 5.90 \times 10^{-2}$) or likely due to the fact the experimental value was not taken at 298.15K (hence removed as an anomaly during analysis). Systematic errors relevant to the entire methodology are likely to have contributed to this as well, which will be explored further in the following sections.

12.2 Uncertainties:

Apparatus	**Uncertainty**	**Smallest value measured**	**Percentage uncertainty**	**Effect on data in context of error**	**Improvements**
Multimeter	±0.001V	0.389V	$\frac{0.001}{0.389} \times 100 = \pm 2.57 \times 10^{-1}\%$	Percentage uncertainties of less than 5% are negligible. These instruments are highly precise.	No improvements are necessary, to decrease systematic error there is not a need to increase the precision of the instruments
Ruler	±0.05cm	3cm	$\frac{0.05}{3} \times 100 = \pm 1.67 \times 10^{0}\%$		
Thermometer	±0.5K	298.15K	$\frac{0.5}{298.15} \times 100 = \pm 1.67 \times 10^{0}\%$		

Measuring Cylinder	50±5cm^3	50cm^3	$\frac{5}{50} \times 100 = \pm 1.00 \times 10^{1}$	This error is greater than 5%, therefore it is significant, it can cause the solutions to not be equimolar, which affects the cell potential due varying moles of ions in E_{cell}. This is a source of random error.	A measuring cylinder with increments of 1 cm^3 rather than 10cm^3 should be used to decrease the uncertainty to ±0.5cm^3 ultimately decreasing the percentage uncertainty to 0.01%

12.3 Systematic Evaluation

Strength	**Benefit to investigation**	
Replacement and removal of the salt bridge	By replacing the salt bridge every set of data, the potassium and nitrate ions were replenished. This was advantageous because these ions were used to reduce positive or negative charges formed in the half cells after redox reactions.	
Replacement of electrodes	Replacing the electrodes allowed for fair testing to take place. It means that every set of trials will have the same surface area for reduction and oxidation to take place, reducing the systematic error for cell potential results	
Controlling the independent variable	The independent variable of temperature was controlled accurately using an electric water bath. This water bath possessed the necessary insulation properties to effectively control the temperature. Furthermore, it decreases the health and safety risk by displaying when the water should not be in contact with the user	
Weakness	**Detriment to investigation**	**Improvement**
Continuous use of the same metal sulphate solution	By using the same metal sulphate solution, the likelihood of non-equimolar solutions increases due to the transfer of ions during the experiment	Using a smaller volume of metal sulphate solution at the same concentration, replaced at every trial
Poor approach to qualitative data collection	Qualitative data was recorded based on touch and was a poor substantiation to the claims made in the analysis	With the use of a magnifying tool, the discrepancies between electrodes at different temperatures can be seen and recorded. Alternatively, the qualitative approach could have been taken by measuring the mass of electrodes before and after being in the galvanic cell
Assumptions of the Q_c	The Q_c is assumed to be constant because trials for each temperature took place in a very short time frame, however this directly goes against scientific theory which explicitly states that Q_c will always tend toward K_c. This affects the calculated equilibrium constant values	Chemical concentration monitors can be made use of to see how the reaction quotient changes over time, this would provide an indication as to the extent in which this assumption can be made.
Varying Electron mobility in electrical wires.	Electron mobility in electrical wires can be affected by the temperature. Since the wires often dipped in and out of the water bath, the temperature of the wires themselves would have changed. This could have caused changes to the cell potential as the resistance to the flow of electrons would have fluctuated	The electrical wires could be elevated by a clamp stand in the air.

13. Conclusion In conclusion the research question ‘How does increasing temperature (298.15±0.50K, 303.15±0.50K, 308.15±0.50K, 313.15±0.50K, 318.15±0.50K) impact cell potential or E_{cell} (V), and hence equilibrium constant (K_c), in the iron-zinc galvanic cell?’ has been answered by finding a negative correlation between increasing temperature and equilibrium constant (K_c), and increasing temperature and cell potential (E_{cell}). It is recognized that this research is only a glimpse of what is required to replace lithium-based batteries and the effectiveness of iron-zinc galvanic cells for use in pacemakers. Though this investigation was not conducted under industry lab conditions, it is accurate to some extent due to the precision of the equipment. However, this could greatly be improved had the relevant changes mentioned in the evaluation been made. One way this investigation could be extended is by finding the relationship between concentration, as opposed to temperature, on cell potential (E_{cell}). This would mean changing the reaction quotient and using the Nernst equation yet again due to non-standard conditions. Another approach to this specific investigation would be to use the Van’t Hoff equation, which also looks at the relationship between temperature and equilibrium constant.

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