CHEM1050 WET labs

General Chemistry II (University of Guelph)

Part 1: Magnesium

First Determination

M*
~~3:33:20~~

Time (min)	Temp. (°C)	Time (min)	Temp. (°C)
3:30:00	21.1	3:39:00	31.9
3:30:30	21.1	3:39:30	31.8
~~3:30:30~~ ~~3:34~~ 3:31:00	21.1	3:40:00	31.7
3:31:30	21.1	3:40:30	31.5
3:32:00	21.1	3:41:00	31.4
3:32:30	21.1	3:41:30	31.1
3:33:00	~~21.1~~ M	3:42:00	31.1
3:33:30	29.0	3:42:30	31.1
3:34:00	29.5	3:43:00	31.0
3:34:30	~~28.6~~ 33.6	3:43:30	31.0
3:35:00	33.7	3:44:00	30.9
3:35:30	33.6	3:44:30	30.9
~~3:36~~ 3:36:00	33.4	3:45:00	30.8
3:36:30	32.8	3:45:30	30.8
3:77:00	32.5	3:46:00	30.6
3:37:30	32.5	3:46:30	30.5
3:38:00	32.1	3:47:00	30.4
3:38:30	32.1	3:47:30	30.3

Second Determination

M*
M*

Time (min)	Temp. (°C)	Time (min)	Temp. (°C)
4:00:00	~~21.1~~ 21.1	4:09:00	32.1
4:00:30	21.1	4:09:30	32.0
4:01:00	21.1	4:10:00	31.9
4:01:30	21.1	4:10:30	31.8
4:02:00	21.1	4:11:00	31.8
4:02:30	21.1	4:11:30	31.8
4:03:00	~~21.1~~ M	4:12:00	31.8
4:03:30	37.0	4:12:30	31.8
4:04:00	36.2	4:13:00	31.7
4:04:30	35.5	4:13:30	31.7
4:05:00	34.8	4:14:00	31.6
4:05:30	34.0	4:14:30	31.5
4:06:00	33.5	4:15:00	31.5
4:06:30	33.0	4:15:30	31.5
4:07:00	32.7	4:16:00	31.4
4:07:30	32.5	4:16:30	31.4
4:08:00	32.4	4:17:00	31.4
4:08:30	32.1	4:17:30	31.4

Part 1	First Determination	Second Determination
Mass of calorimeter + HCl	199.24g	200.05g
Mass of calorimeter	139.38g	140.06g
Mass of hydrochloric acid	59.86g	59.99g
Mass of vial + magnesium	15.1573g	15.1757g
Mass of vial	14.9741g	14.9899g
Mass of magnesium	0.1832g	0.1858g
Moles of magnesium	$n_{Mg} = \frac{1 mol Mg}{24.305 g Mg} \times 0.1832 g Mg$ $= 0.007537543$ mol $= 7.538 \times 10^{-3}$ mol 7.538×10^{-3} mol ✓	$n_{Mg} = \frac{1 mol Mg}{24.305 g Mg} \times 0.1858 g Mg$ $= 0.007644517$ mol Mg $= 7.644$ mol Mg 7.644×10^{-3} mol
Heat capacity of final solution $C_s = 3.92\ JK^{-1}g^{-1}$	$C_p = C_s \times$ mass used of 1M HCl $C_p = 3.92\ JK^{-1}g^{-1} \times 59.86 g$ $= 234.6512\ JK^{-1}$ $= 235\ JK^{-1}$ ✓ $235\ JK^{-1}$	$C_p = 3.92\ JK^{-1}g^{-1} \times 59.99 g$ $= 235.1608\ JK^{-1}$ $235\ JK^{-1}$

ΔT (from graph) 10.4°C 12.2°C

Heat flow, q_{rxn}

$$Mg_{(s)} + 2H^{+}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + H_{2(g)}$$

$$q_{rxn} = -C_p \times \Delta T$$

$q_{rxn} = -(235\ J\ K^{-1})(10.4\ K)$

$= -2444\ J$

$= -2.444\ kJ$

$= -2.44\ kJ$

$q_{rxn} = -(235\ J\ K^{-1})(12.2\ K)$

$= -2867\ J$

$= -2.867\ kJ$

$= -2.87\ kJ$

✓ −2.44 kJ −2.87 kJ

ΔH_f° of Mg^{2+}(aq)

$$q_{rxn} \div n_{Mg} = \Delta H_f^\circ \text{ per mol of } Mg^{2+}_{(aq)}$$

$\Delta H = \dfrac{-2.44\ kJ}{7.538 \times 10^{-3}\ mol}$

$= -324\ kJ/mol$

$\Delta H = \dfrac{-2.87\ kJ}{7.644 \times 10^{-3}\ mol}$

$= -374\ kJ/mol$

✓ −324 kJ/mol −374 kJ/mol

Average ΔH_f° of Mg^{2+}(aq)

$= \dfrac{[(-324)+(-374)]\ kJ/mol}{2}$

$= -349\ kJ/mol$

−349 kJ/mol

Part 2: Magnesium Oxide

First Determination

M*

Time (min)	Temp. (°C)	Time (min)	Temp. (°C)
4:30:00	21.1	4:39:00	28.9
4:30:30	21.1	4:39:30	28.9
4:31:00	21.1	4:40:00	28.9
4:31:30	21.1	4:40:30	28.9
4:32:00	21.1	4:41:00	28.8
4:32:30	21.1	4:41:30	28.8
4:33:00	~~22.9~~ ~~21.1~~ M	4:42:00	28.6
4:33:30	27.9	4:42:30	28.1
4:34:00	28.9	4:43:00	28.0
4:34:30	29.0	4:43:30	28.0
4:35:00	29.0	4:44:00	28.0
4:35:30	28.9	4:44:30	27.9
4:36:00	28.9	4:45:00	27.9
4:36:30	29.0	4:45:30	27.9
4:37:00	29.0	4:46:00	27.9
4:37:30	29.1	4:46:30	27.8
4:38:00	28.9	4:47:00	27.8
4:38:30	28.8	4:47:30	27.8

Second Determination

M*

Time (min)	Temp. (°C)	Time (min)	Temp. (°C)
5:00:00	21.1	5:09:00	28.5
5:00:30	21.1	5:09:30	28.2
5:01:00	21.1	5:10:00	28.2
5:01:30	21.1	5:10:30	28.2
5:02:00	21.1	5:11:00	28.2
5:02:30	21.1	5:11:30	28.1
5:03:00	~~21.1~~ M	5:12:00	28.0
5:03:30	27.0	5:12:30	28.0
5:04:00	28.9	5:13:00	28.0
5:04:30	29.0	5:13:30	28.0
5:05:00	29.1	5:14:00	28.0
5:05:30	29.0	5:14:30	27.9
5:06:00	29.0	5:15:00	27.9
5:06:30	29.0	5:15:30	27.9
5:07:00	29.0	5:16:00	27.9
5:07:30	29.0	5:16:30	27.9
5:08:00	28.9	5:17:00	27.9
5:08:30	28.8	5:17:30	27.9

Part 2	First Determination	Second Determination
Mass of calorimeter + acid (HCl)	200.72g	200.56g
Mass of calorimeter	140.13 g	140.28g
Mass of hydrochloric acid	60.59g	60.28g
Mass of vial + MgO	14.1816g	14.1839g
Mass of vial	13.4821g	13.4843g
Mass of MgO	0.6995g	0.6996g
Moles of MgO $MM_{MgO} = 40.304 g/mol$	$n_{MgO} = \frac{1 mol MgO}{40.304 g MgO} \times 0.6995 g MgO$ $= 0.017355 mol MgO$ $= 1.736 \times 10^{-2} mol MgO$ ✓	$n_{MgO} = \frac{1 mol MgO}{40.304 g MgO} \times 0.6996 g MgO$ $= 0.017358 mol MgO$ $= 1.736 \times 10^{-2} mol MgO$
	1.736×10^{-2} mol MgO	1.736×10^{-2} mol MgO
Heat capacity of final solution $C_s = 3.92 J K^{-1} g^{-1}$	$C_p = C_s \times$ mass 1M HCl used $C_p = 3.92 J K^{-1} g^{-1} \times 60.59 g$ $= 238 J K^{-1}$ ✓	$C_p = 3.92 J K^{-1} g^{-1} \times 60.28 g$ $= 236 J K^{-1}$
	238 JK^{-1}	236 JK^{-1}
ΔT (from graph)	8.6°C	8.5°C

Heat flow, q_{rxn}

$q_{rxn} = -C_p \times \Delta T$

$= (-238\ JK^{-1})(8.6\ K)$ $\quad$ $= -(236\ JK^{-1})(8.5\ K)$

$= -2046.8\ J$ $\quad$ $= 2006\ J$

$-2.05\ kJ$ $\quad$ $-2.01\ kJ$

sig fig (-1)

ΔH of reaction per mole of MgO

$\Delta H = \frac{-2.05\ kJ}{1.736 \times 10^{-2}\ mol}$ $\quad$ $\Delta H = \frac{-2.01\ kJ}{1.736 \times 10^{-2}\ mol}$

$= -118\ kJ/mol$ $\quad$ $= -115.78\ kJ/mol$

$= -116\ kJ/mol$

$-118\ kJ/mol$ $\quad$ $-116\ kJ/mol$

Average ΔH of reaction of MgO with HCl $-117\ kJ/mol$

Enthalpy of Formation of Magnesium Oxide

Use the average values from your experimental work in your calculation of ΔH_f° of MgO.

$$\Delta H = \Delta H_f^\circ[Mg^{2+}_{(aq)}] + \Delta H_f^\circ[H_2O_{(l)}] - \Delta H_f^\circ[MgO_{(s)}] - 2\Delta H_f^\circ(H^+_{(aq)}]$$

$$(-117\ kJ/mol) = (-349\ kJ/mol) + (-285.8\ kJ/mol) - \Delta H_f^\circ[MgO_{(s)}] - 2(0\ kJ/mol)$$

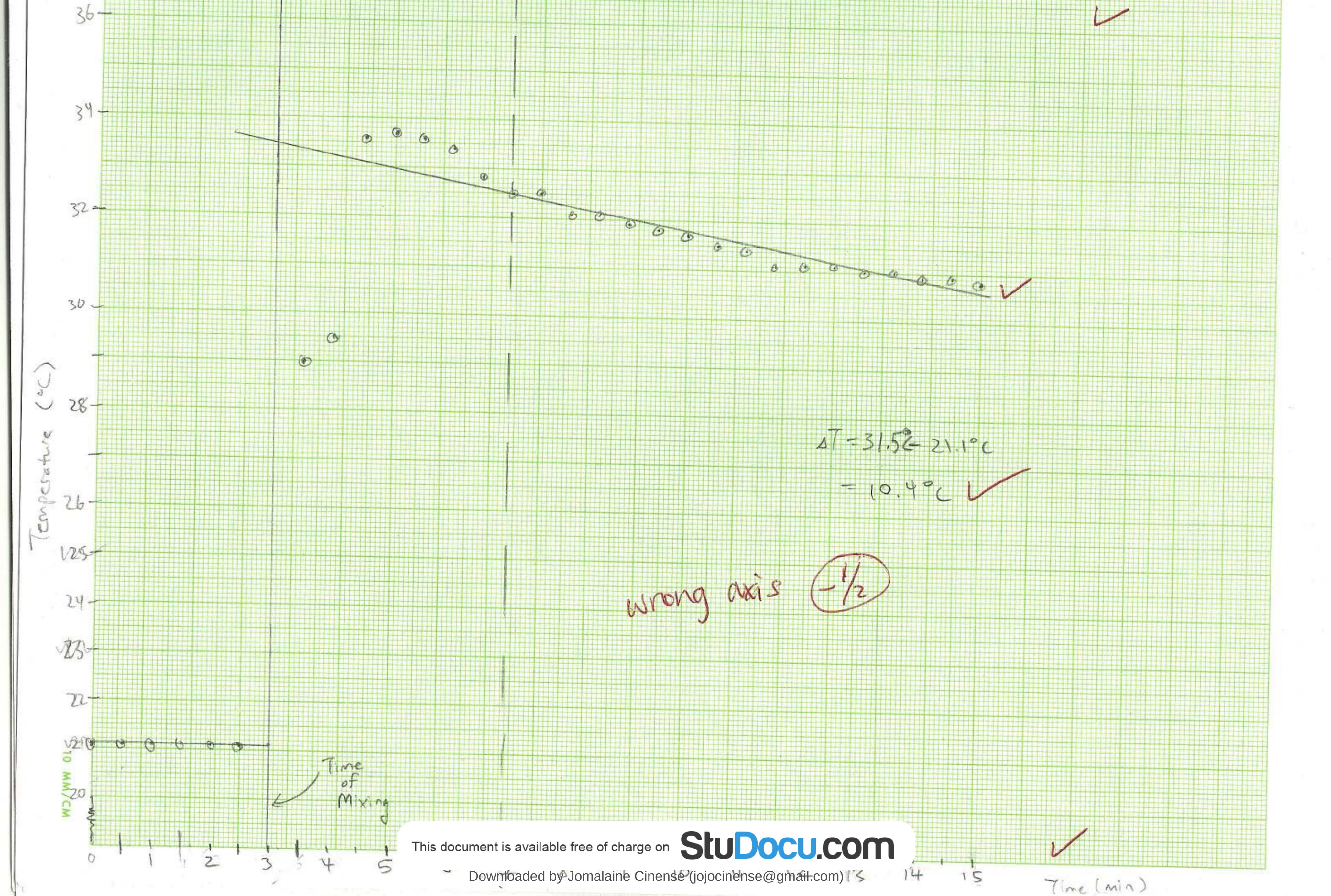
$$\Delta H_f^\circ(MgO_{(s)}] = -349\ kJ/mol - 285.8\ kJ/mol + 117\ kJ/mol$$

$$= -517.8\ kJ/mol$$

sig fig (-1)

$15\ min \times \frac{60\ sec}{1\ min} =$

Enthalpy of Formation of the 1st Determination of 0.1832g Magnesium with 59.86g 1M Hydrochloric Acid

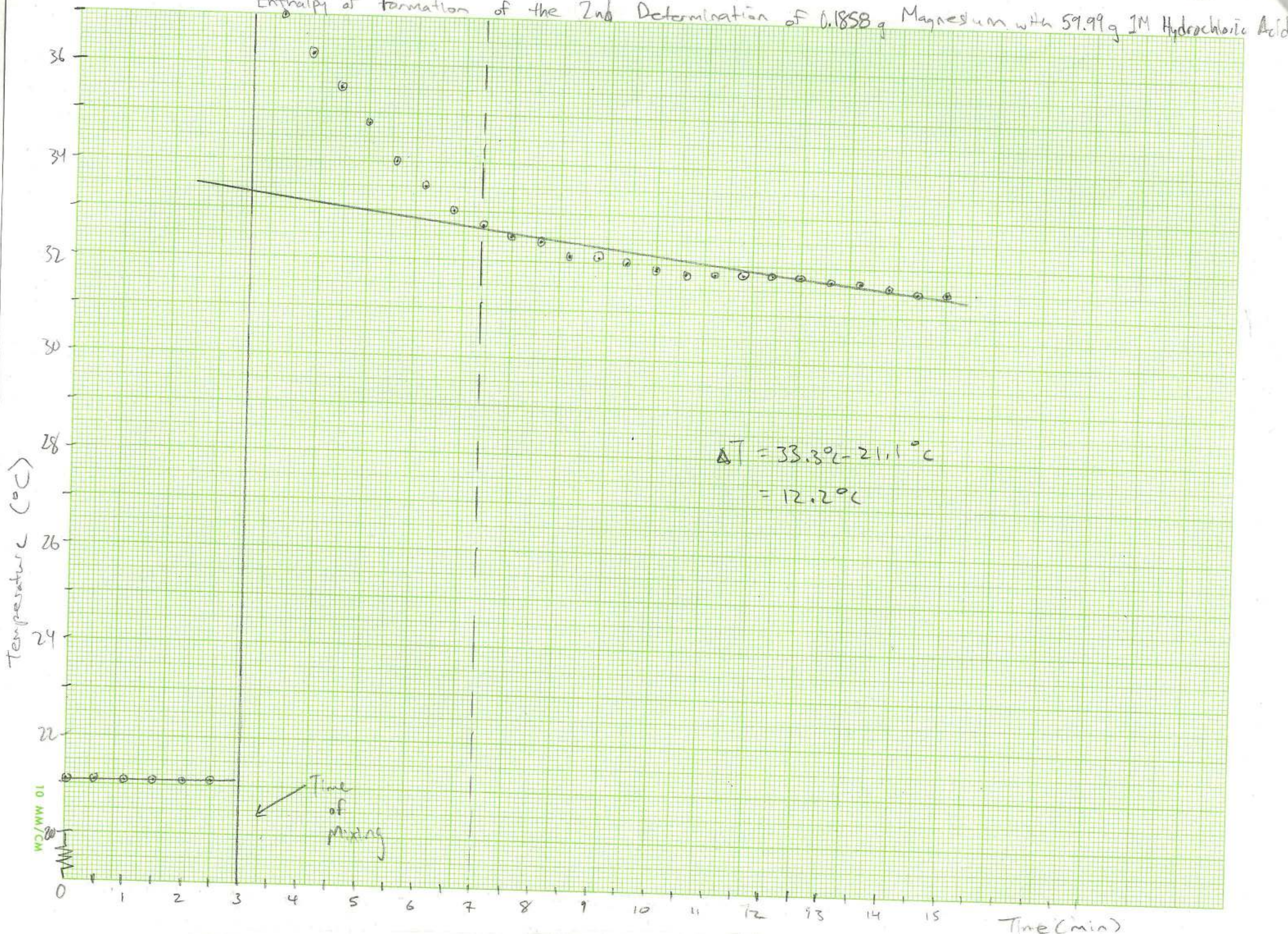


ΔT = 31.5°C − 21.1°C

= 10.4°C ✓

wrong axis −1/2

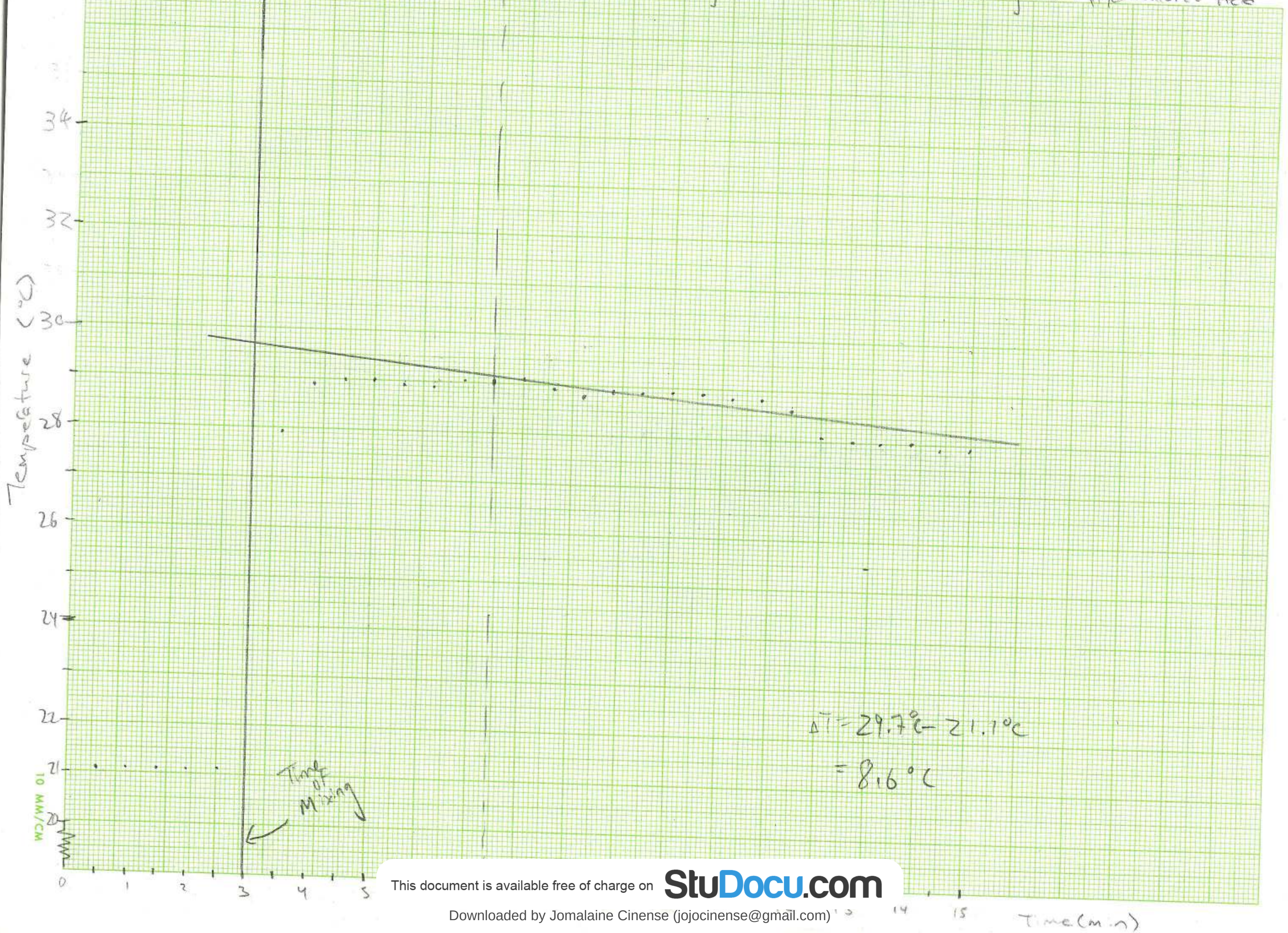
Enthalpy of Formation of the 2nd Determination of 0.1858 g Magnesium with 59.99 g 1M Hydrochloric Acid



ΔT = 33.3°C − 21.1°C
= 12.2°C

1D

Enthalpy of Formation of the 1st Determination of 0.6995g Magnesium Oxide with 60.59g 1M Hydrochloric Acid

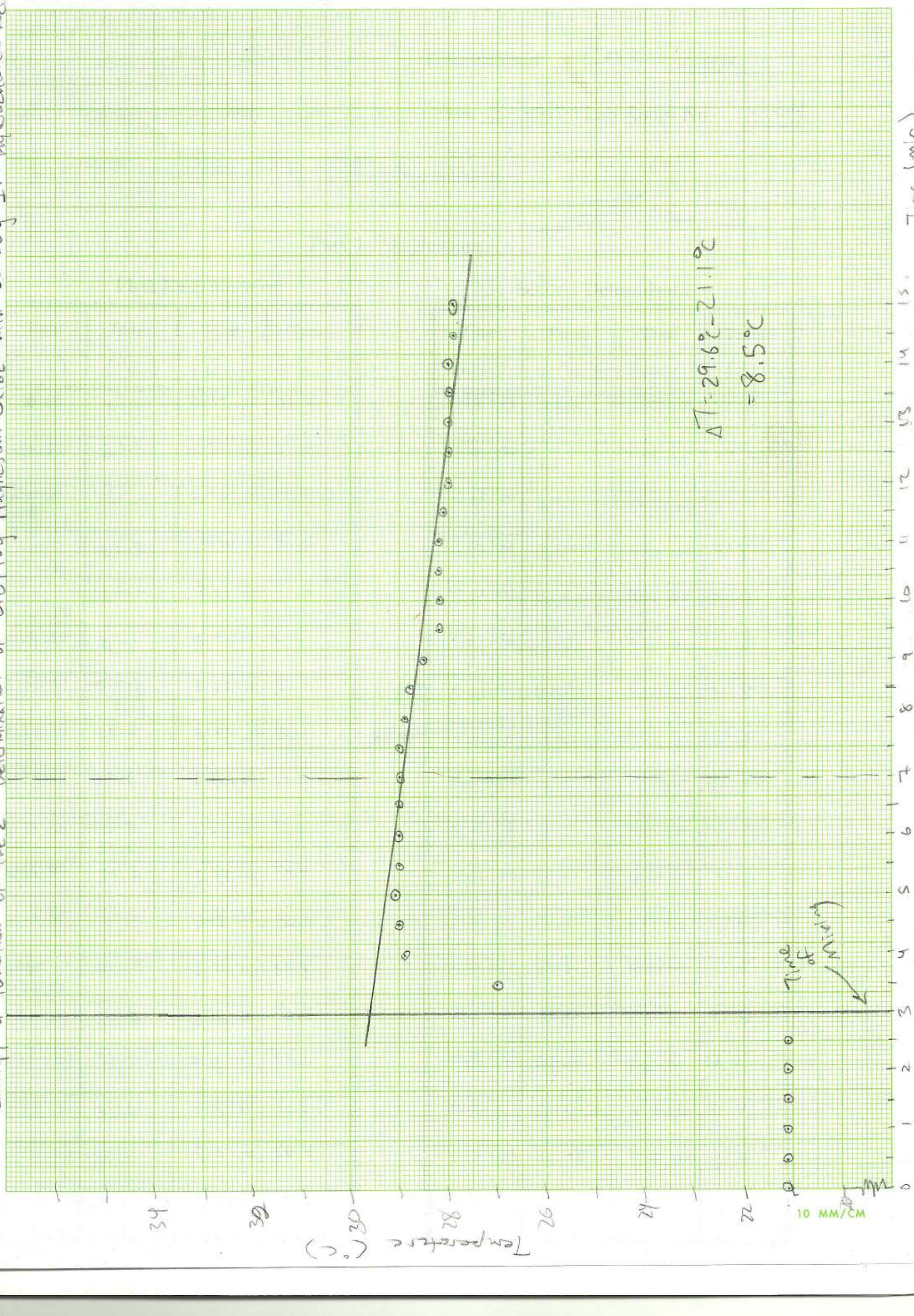


ΔT = 29.7°C − 21.1°C

= 8.6°C

1)

Enthalpy of Formation of the 2nd Determination of 0.6996 g Magnesium Oxide with 60.28 g 1 M Hydrochloric Acid



ΔT = 29.6°C − 21.1°C
= 8.5°C

Concentration of standard $Na_2S_2O_3$ solution 0.01066 M

Temperature of the Equilibrium Mixture 20.5 °C

10/10

Part 1 Aqueous layer

Volume of Aqueous Aliquots:	1	2	3
final buret reading	17.52 mL	27.59 mL	37.71 mL
initial buret reading	~~8.52~~ 7.48 mL	17.52 mL	27.59 mL
volume of aliquot	10.04 mL	10.07 mL	10.12 mL
Titration with Sodium Thiosulfate final buret reading	~~1.79~~ 32.59 mL	~~32.25~~ 34.55 mL	35.65 mL
initial buret reading	1.79 mL	1.61 mL	2.84 mL
volume of thiosulfate titer	~~32.59~~ 30.80 mL	~~32.25~~ 32.94 mL	32.81 mL
Total concentration of reducible iodine in aqueous phase	1.635×10^{-2} mol/L	1.743×10^{-2} mol/L (BEST)	1.730×10^{-2} mol/L (BEST)

Average of "best two" 1.736×10^{-2} mol/L
(Label the best two titrations)

$$I_{2(aq)} + 2S_2O_3^{2-} \rightarrow 2I^-_{(aq)} + S_4O_6^{2-}{}_{(aq)}$$

Trial 1:

$$n_{Na_2S_2O_3} = (0.01066 \text{ mol/L})(0.03080 \text{ L})$$
$$= 3.283 \times 10^{-4} \text{ mol}$$

$$n_{I_2} = 3.283 \times 10^{-4} \text{ mol } S_2O_3^{2-} \times \frac{1 \text{ mol } I_2}{2 \text{ mol } S_2O_3^{2-}}$$
$$= 1.6416 \times 10^{-4} \text{ mol } I_2$$
$$= 1.642 \times 10^{-4} \text{ mol } I_2$$

$$\therefore [I_2] = \frac{1.6416 \times 10^{-4} \text{ mol}}{0.01004 \text{ L}}$$
$$= 1.635 \times 10^{-2} \text{ mol/L}$$

more space on next page

$n_{Na_2S_2O_3} = (0.01066\ mol/L)(0.03294 L)$

$= 0.0003511\ mol\ Na_2S_2O_3$

$= 3.511 \times 10^{-4}\ mol\ Na_2S_2O_3$

$n_{I_2} = 3.511 \times 10^{-4}\ mol\ Na_2S_2O_3 \times \frac{1\ mol\ I_2}{2\ mol\ Na_2S_2O_3}$

$= 1.7557 \times 10^{-4}\ mol\ I_2$

$= 1.756 \times 10^{-4}\ mol\ I_2$

$[I_2] = \frac{1.7557\ mol\ I_2}{0.01007\ L}$

$= 1.743 \times 10^{-2}\ mol/L$

$n_{Na_2S_2O_3} = (0.01066\ mol/L)(0.03281\ L)$

$= 3.5007 \times 10^{-4}\ mol\ Na_2S_2O_3$

$= 3.501 \times 10^{-4}\ mol\ Na_2S_2O_3$

$n_{I_2} = 3.501 \times 10^{-4}\ mol\ Na_2S_2O_3 \times \frac{1\ mol\ I_2}{2\ mol\ Na_2S_2O_3}$

$= 1.750 \times 10^{-4}\ mol\ I_2$

$[I_2] = \frac{1.750 \times 10^{-4}\ mol\ I_2}{0.01012\ L}$

$= \underline{1.7297 \times 10^{-2}\ mol/L}$

$= 1.730 \times 10^{-2}\ mol/L$

Part II CH_2Cl_2 Layer

Volume of CH_2Cl_2 Aliquots:	1	2	3
final buret reading	33.71 mL	40.68 mL	47.72 mL
initial buret reading	26.69 mL	33.71 mL	40.68 mL
volume of aliquot	7.02 mL	6.97 mL	7.04 mL
Titration with Sodium Thiosulfate			
final buret reading	31.49 mL	28.59 mL	28.71 mL
initial buret reading	5.51 mL	2.08 mL	1.98 mL
volume of thiosulfate titer	26.28 mL	26.51 mL	26.73 mL
Concentration of iodine in CH_2Cl_2	2.00×10^{-2} mol/L	2.03×10^{-2} mol/L (BEST)	2.02×10^{-2} mol/L (BEST)

Average of "best two" (Label the best two titrations) 2.02×10^{-2} mol/L

$I_{2(aq)} \rightleftharpoons I_{2(CH_2Cl_2)}$ $I_{2(aq)} + 2S_2O_3{}^{2-}{}_{(aq)} \longrightarrow 2I^-{}_{(aq)} + S_4O_6{}^{2-}$

Trial 1:

$$n_{I_2} = 0.02688\,L\ S_2O_3^{2-} \times \frac{0.01066\,mol\ S_2O_3^{2-}}{1\,L\ S_2O_3^{2-}} \times \frac{1\,mol\ I_2}{2\,mol\ S_2O_3^{2-}}$$

$= 1.401 \times 10^{-4}$ mol I_2

$n_{I_2} = n_{I_2(CH_2Cl_2)}$

$$[I_2]_{CH_2Cl_2} = \frac{1.401 \times 10^{-4}\,mol}{0.00702\,L}$$

$= 1.995 \times 10^{-2}$ mol/L

$= 2.00 \times 10^{-2}$ mol/L

Trial 2:

$$n_{I_2} = 0.02651\,L\ S_2O_3{}^{2-} \times \frac{0.01066\,mol\ S_2O_3^{2-}}{1\,L\ S_2O_3{}^{2-}} \times \frac{1\,mol\ I_2}{2\,mol\ S_2O_3^{2-}}$$

$= 1.413 \times 10^{-4}$ I_2 ; $n_{I_2} = n_{I_2(CH_2Cl_2)}$

$$[I_2]_{CH_2Cl_2} = \frac{1.413 \times 10^{-4}\,mol\ I_2}{0.00697\,L}$$

$= 2.03 \times 10^{-2}$ mol/L

Trial 3:

$$n_{I_2} = 0.02673\,L\ S_2O_3^{2-} \times \frac{0.01066\,mol\ S_2O_3^{2-}}{1\,L\ S_2O_3^{2-}} \times \frac{1\,mol\ I_2}{2\,mol\ S_2O_3^{2-}}$$

$= 1.425 \times 10^{-4}$ mol I_2

$n_{I_2} = n_{I_2(CH_2Cl_2)}$

$$[I_2]_{CH_2Cl_2} = \frac{1.425 \times 10^{-4}\,mol}{0.00704\,L}$$

$= 2.02 \times 10^{-2}$ mol/L

Concentration of I_2 in water

$[I_2]_{CH_2Cl_2} = 2.02 \times 10^{-2}$ mol/L $K = 176$

$$K = \frac{[I_2]_{CH_2Cl_2}}{[I_2]_{H_2O}}$$

$$[I_2]_{H_2O} = \frac{[I_2]_{CH_2Cl_2}}{K} = \frac{2.02 \times 10^{-2} \text{ mol/L}}{176} = 1.15 \times 10^{-4} \text{ mol/L}$$

1.15×10^{-4} mol/L

Concentration of I_3^- in water

$[I_3^-] = [I_2] - [I_2]_{H_2O}$

$= 1.736 \times 10^{-2}$ mol/L $- 1.15 \times 10^{-4}$ mol/L

$= 1.72 \times 10^{-2}$ mol/L

1.72×10^{-2} mol/L

Concentration of I^- in water

$[K^+]_{(aq)} = [I^-]_{(aq)} + [I_3^-]_{(aq)} = 0.200$ mol/L

$[I^-]_{(aq)} = 0.200$ mol/L $- [I_3^-]_{(aq)}$

$= 0.200$ mol/L $- 1.72 \times 10^{-2}$ mol/L

$= 0.182755$ mol/L

$= 1.83 \times 10^{-1}$ mol/L

1.83×10^{-1} mol/L

Value of K_c in water, at Temperature 20.5°C °C

$$K_c = \frac{[I_3^-]_{(aq)}}{[I_2]_{(aq)}[I^-]_{(aq)}}$$

$$= \frac{(1.72 \times 10^{-2})}{(1.15 \times 10^{-4})(1.83 \times 10^{-1})}$$

$= 817$

Part A Determination of the Order of Reducing Strengths of Zn, Pb, Cu and Fe

1. On the basis of your experimental observations on Report Pages 2 and 3, arrange the four metals (not ions) in order of decreasing reducing strength. Give your reasoning, which must follow from the results alone. The results usually make it impossible to arrange all four metals unambiguously; in that case order as many as you can, and then show/discuss clearly all the ambiguities and contradictions which concern the metals you cannot order. Show all the positions where these metals "might" fall according to your data. Hint: the **number** of reductions that occurred, the **speed**, and the **extent** of a reaction are irrelevant. Do not discuss reactions that do **not** occur.

   Do not use any form of the word "oxidize". You may use any form of the word "reduce".

The order of decreasing reducing strength based upon the obtained experimental results is:

$Zn_{(s)}$ is the strongest reducing agent because it reacted to $Cu^{2+}_{(aq)}$, $Pb^{2+}_{(aq)}$ and $Fe^{2+}_{(aq)}$, where all three were identified as having a colour change.

$Pb_{(s)}$ is the second strongest reducing agent since it reacted with both $Cu^{2+}_{(aq)}$ and $Fe^{2+}_{(aq)}$. $Zn_{(s)}$ reacted with $Cu^{2+}_{(aq)}$, but $Cu_{(s)}$ did not react back with $Zn^{2+}_{(aq)}$ showing that $Zn_{(s)}$ is the strongest reducing agent; this shows that $Cu_{(s)}$ is also the weaker reducing agent. This also applies to $Pb_{(s)}$, since $Pb_{(s)}$ reacted with $Cu^{2+}_{(aq)}$, but not vice versa, it shows that the reaction is spontaneous in the reverse direction. ~~Due to a~~ Since a colour change was observed, it means $Zn_{(s)}$ reacted with $Fe^{2+}_{(aq)}$. Since ~~$Fe^{2+}_{(aq)}$~~ $Fe_{(s)}$ did not react with $Zn^{2+}_{(aq)}$ however, it shows that $Fe_{(s)}$ is a weaker reducing agent than $Zn_{(s)}$. Lastly, since $Fe_{(s)}$ reacted with $Cu^{2+}_{(aq)}$ and ~~Fe~~ $Cu_{(s)}$ also reacted with $Fe^{2+}_{(aq)}$, it is uncertain which of the two is the stronger reducing agent.

incomplete order

−1

2. In each square, answer the following questions:
   - a) Do you think that a reaction occurred? Check YES or NO
   - b) If a reaction occurred, write your observations.
   - c) If a reaction occurred, write a balanced net ionic equation.

Ions Metals	Cu^{2+}	Zn^{2+}
Cu		Yes ● No ☑
Zn	Yes ☑ No ● - became black in colour - surface became rough + thick $Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)}$	
Pb	Yes ☑ No ● - ~~purple~~ tip turned purple in colour (overall darker) $Pb_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cu_{(s)} + Pb^{2+}_{(aq)}$	Yes ● No ☑
Fe	Yes ☑ No ● - tip turned reddish / brown in colour $Fe_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cu_{(s)} + Fe^{2+}_{(aq)}$	Yes ● No ☑

Ions Metals	Pb^{2+}	Fe^{2+}
Cu	Yes ● No ☑ ~~white precipitate~~	Yes ☑ No ● - surface of metal became lighter in colour $Cu_{(s)} + Fe^{2+}_{(aq)} \rightarrow Fe_{(s)} + Cu^{2+}_{(aq)}$
Zn	Yes ☑ No ● - tip turns dull black/greyish in colour - formation of white crystal-like substances $Zn_{(s)} + Pb^{2+}_{(aq)} \rightarrow Pb_{(s)} + Zn^{2+}_{(aq)}$	Yes ☑ No ● - surface more dark than before $Zn_{(s)} + Fe^{2+}_{(aq)} \rightarrow Fe_{(s)} + Zn^{2+}_{(aq)}$
Pb		Yes ☑ No ● - surface turned blue in colour $Pb_{(s)} + Fe^{2+}_{(aq)} \rightarrow Pb^{2+}_{(aq)} + Fe_{(s)}$
Fe	Yes ● No ☑	

Part B Measurement of Cell Potentials

1. Because the Fe^{2+} concentration is not 0.5 M, the experimental voltages can be expected to deviate from the theoretical voltages in all cases involving the Fe^{2+} half-cell.

Cell	Experimental Voltages and Average	Name of Anode	Name of Cathode	Theoretical Voltage (Assume 0.5 M)
R Zn-Cu	① 1.073 V ② 1.074 V Average: =1.074V	Zn	Cu	1.10V
R Cu-Pb	① 0.460 V ② 0.460 V Average: = 0.460V	Pb	Cu	0.47V
R Cu-Fe	① 0.510 V ② 0.518 V Average: =0.514V	Fe	Cu	0.78V
R Zn-Pb	① 0.612 V ② 0.612 V Average: =0.612V	Zn	Pb	0.63V
r Fe-Zn	① 0.540V ② 0.533 V Average: =0.537V	Zn	Fe	0.32V
r Fe-Pb	① 0.065 V ② 0.066 V Average: = 0.066V	Fe	Pb	0.31V

2. For the Cu-Pb cell, write the equation for the

Anode Reaction $Pb_{(s)} \rightarrow Pb^{2+}_{(aq)} + 2e^-$

Cathode Reaction $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

3. On the basis of the theoretical voltages, write the order for the reducing strength of the four metals, in order of decreasing reducing strength.

The order for the 4 metals in decreasing reducing strength is:

$Zn_{(s)} > Fe_{(s)} > Pb_{(s)} > Cu_{(s)}$

4. One notices greater than average discrepancies between the theoretical and experimental voltages each time the Fe/Fe^{2+} half-cell is involved. This indicates that the concentration of the Fe^{+2} solution was not 0.5 mol/L. On the basis of your measurements from Part B and using the Nernst equation decide whether the Fe^{2+} solution concentration was greater or smaller than 0.5 mol/L. **Do no calculations.** Clearly show your reasoning algebraically with references to specific cells. Use **all** evidence available.

The Fe^{2+} solution concentration was greater than 0.5 mol/L.

For Part B: $Cu^{2+}_{(aq)} + Fe_{(s)} \longrightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$

$Pb^{2+}_{(aq)} + Fe_{(s)} \longrightarrow Fe^{2+}_{(aq)} + Pb_{(s)}$ ($Fe_{(s)}$ is the reducing agent in both equations)

for $Q = \frac{[Fe^{2+}]}{[Cu^{2+}]} = \frac{[Fe^{2+}]}{[Pb^{2+}]} = \frac{[Fe^{2+}]}{0.5}$ and E° is positive

Using the Nernst equation:

$$E = E^\circ - \frac{RT}{nF}\ln Q$$

$= E^\circ - 0.0296 \log\left(\frac{[Fe^{2+}]}{0.5}\right)$ where $[Fe^{2+}] > 0.5$ mol/L giving $\log Q > 0$

∴ it can be concluded that a positive E° subtracting a small positive number decreases E. Referring to Part B, the E values for the experiment for Cu-Fe cell and Fe-Pb cell was lower in comparison to the theoretical value of the cell.

For $Fe^{2+}_{(aq)} + Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)} + Fe_{(s)}$

$Q = \frac{[Zn^{2+}]}{[Fe^{2+}]} = \frac{0.5}{[Fe^{2+}]}$ and E° is positive

Using the Nernst Equation:

$$E = E^\circ - \frac{RT}{nF}\ln Q$$

$= E^\circ - 0.0296 \log\left(\frac{0.5}{[Fe^{2+}]}\right)$ where $[Fe^{2+}] > 0.5$ mol/L giving $\log Q < 0$

$\therefore E = E^\circ + 0.0296 \log Q$

This on the other hand gives an increase in E, which can be seen from the experiment value for Fe-Zn in comparison to the theoretical value.

Part C Influence of Concentration on Cell Voltage

Measured Temperature of Zn–Cu cell in **Part C** 22.2 °C

Measured Voltage of the Zn–Cu cell in **Part C**, with 0.005 M Cu^{2+} 1.045 V

(Measured Voltage of the Zn–Cu cell in **Part B**, at 0.5 M) 1.074 V

1. Describe what happened to the measured voltage when the cathode solution was diluted.

The measured voltage decreased.

2. What do you expect would have happened if the anode solution had been diluted, instead of the cathode solution?

The measured voltage would be expected to increase.

3. Calculate the theoretical voltage of the Zn – Cu cell in **Part C**, using standard reduction potentials and the measured temperature.

$$E^\circ = E^\circ(Cu^{2+}/Cu) - E^\circ(Zn^{2+}/Zn)$$
$$= 0.34V - (-0.76V)$$
$$= 1.10V$$

$$E = E^\circ - \frac{RT}{nF}\ln Q$$
$$= E^\circ - \frac{RT}{nF}\ln\frac{[Zn^{2+}]}{[Cu^{2+}]}$$

$$= 1.10V - \frac{(8.314\,JK^{-1}mol^{-1})(295.2\,K)}{2(96485\,C\,mol^{-1})}\ln\left(\frac{0.5}{0.005}\right)$$
$$= 1.10V - 0.05857\,V$$
$$= 1.04\ V$$

Part A. Dependence of Reaction Rate on Concentration

	1 Time Seconds	2 Temp. °C	3 Initial $[I^-]$ $(mol\ L^{-1})$	4 Initial $[S_2O_8^{2-}]$ $(mol\ L^{-1})$	5 Rate $(mol\ L^{-1}s^{-1})$	6 Relative Rate
1	87s ~~87.1~~	22.0 22.2 22.1	4.00×10^{-2} mol/L	4.00×10^{-2} mol/L	5.75×10^{-6} mol $L^{-1}s^{-1}$	2
2	~~46.7~~ 47s	21.9 21.9 21.9	8.00×10^{-2} mol/L	4.00×10^{-2} mol/L	1.06×10^{-5} mol $L^{-1}s^{-1}$	4
3	~~3~~ 195s ~~195.2~~	21.9 21.5 21.7	2.00×10^{-2} mol/L	4.00×10^{-2} mol/L	2.56×10^{-6} mol $L^{-1}s^{-1}$	1
4	148s ~~147.7~~ a bit large	21.4 21.3 21.4	8.00×10^{-2} mol/L	2.00×10^{-2} mol/L	3.38×10^{-6} mol $L^{-1}s^{-1}$	1
5	169s ~~168.8~~	22.9 22.7 22.8	8.00×10^{-2} mol/L	1.00×10^{-2} mol/L	2.96×10^{-6} mol $L^{-1}s^{-1}$	1

More space for calculations on next page.

umn 3: Calculation of Initial $[I^-]$

$$n_1 = n_2 \Rightarrow (cV)_1 = (cV)_2 \quad \therefore \quad c_2 = \frac{(cV)_1}{V_2}$$

where 1 refers to before mixing
2 refers to after mixing

$$V_2 = 50.0\,mL = 50.0\times10^{-3}\,L$$

Reaction 1: $[I^-] = \dfrac{(0.200\,mol/L)(10.0\times10^{-3}\,L)}{50.0\times10^{-3}\,L} = 4.00\times10^{-2}\,mol/L$

Reaction 2, 4, 5: $[I^-] = \dfrac{(0.200\,mol/L)(20.0\times10^{-3}\,L)}{50.0\times10^{-3}\,L} = 8.00\times10^{-2}\,mol/L$

Reaction 3: $[I^-] = \dfrac{(0.200\,mol/L)(5.00\times10^{-3}\,L)}{50.0\times10^{-3}\,L} = 2.00\times10^{-2}\,mol/L$

Column 4: Calculation of Initial $[S_2O_8^{2-}]$

Reaction 1, 2, 3: $[S_2O_8^{2-}] = \dfrac{(0.100\,mol/L)(20.0\times10^{-3}\,L)}{50.0\times10^{-3}\,L} = 4.00\times10^{-2}\,mol/L$

Reaction 4: $[S_2O_8^{2-}] = \dfrac{(0.100\,mol/L)(10.0\times10^{-3}\,L)}{50.0\times10^{-3}\,L} = 2.00\times10^{-2}\,mol/L$

Reaction 5: $[S_2O_8^{2-}] = \dfrac{(0.100\,mol/L)(5.00\times10^{-3}\,L)}{50.0\times10^{-3}\,L} = 1.00\times10^{-2}\,mol/L$

Column 5: Calculating Rate:

$$rate = \frac{\Delta[S_2O_8^{2-}]}{\Delta t} \qquad \Delta[S_2O_8^{2-}] = \frac{2.50\times10^{-5}\,mol}{50.0\times10^{-3}\,L} = 5.00\times10^{-4}\,mol/L$$

$rate_1 = \dfrac{5.00\times10^{-4}\,mol/L}{87\,s} = 5.75\times10^{-6}\,mol\cdot L^{-1}\cdot s^{-1}$

$rate_2 = \dfrac{5.00\times10^{-4}\,mol/L}{47\,s} = 1.06\times10^{-5}\,mol\cdot L^{-1}\cdot s^{-1}$

$rate_3 = \dfrac{5.00\times10^{-4}\,mol/L}{195\,s} = 2.56\times10^{-6}\,mol\cdot L^{-1}\cdot s^{-1}$

$rate_4 = \dfrac{5.00\times10^{-4}\,mol/L}{148\,s} = 3.38\times10^{-6}\,mol\cdot L^{-1}\cdot s^{-1}$

$rate_5 = \dfrac{5.00\times10^{-4}\,mol/L}{169\,s} = 2.96\times10^{-6}\,mol\cdot L^{-1}\,s^{-1}$

Column 6: Calculating Relative Rate of Reaction:

∴ $rate_3$ has the smallest rate

$rel.rate_1 = \dfrac{rate_1}{rate_5} = \dfrac{5.75\times10^{-6}}{2.56\times10^{-6}} = 2$

$rel.rate_2 = \dfrac{rate_2}{rate_3} = \dfrac{1.06\times10^{-5}}{2.56\times10^{-6}} = 4$

$rel.rate_3 = \dfrac{rate_3}{rate_3} = 1$

$rel.rate_4 = \dfrac{rate_4}{rate_3} = \dfrac{3.38\times10^{-6}}{2.56\times10^{-6}} = 1$

$rel.rate_5 = \dfrac{rate_5}{rate_3} = \dfrac{2.96\times10^{-6}}{2.56\times10^{-6}} = 1$

$rate = k[S_2O_8^{2-}]^m[I^-]^n$

1. Using your data from Part A, what can you conclude about the dependence of the rate on $[I^-]$? (eg., the rate is proportional to $[I^-]$, or to $[I^-]^2$, etc.). Explain fully, with numerical support and reference to all relevant solutions.

From Part A, it can be concluded that the rate of the reaction is proportional to $[I^-]$. Looking at reaction 2 to 1, it can be seen that $[I^-]$ doubles, with reaction 2 having a concentration of 8.00×10^{-2} mol/L and reaction 1 of 4.00×10^{-2} mol/L. The rate also doubles, with 5.75×10^{-6} $mol \cdot L^{-1} \cdot s^{-1}$ coming close to ~ 1.06×10^{-5} $mol \cdot L^{-1} \cdot s^{-1}$. This can also be seen with reaction 1 to 3 with $[I^-]$ doubling as well as the rate. Looking at reaction 2 to 3, $[I^-]$ increases by a factor of four, with the rate also increasing by approximately four. This is observed to be true when $[S_2O_8^{2-}]$ is constant. Thus the reaction is first order in $[I^-]$, with $n = 1$. ✓

2. How does the rate depend on $[S_2O_8^{2-}]$? (Answer as in Question 1 above.)

From Part A, it can be concluded that the rate is proportional to $[S_2O_8^{2-}]$. By examining reaction 2 to 4, $[S_2O_8^{2-}]$ doubles with the rate doubling roughly as well. From reaction 2 to 5, $[S_2O_8^{2-}]$ increases by a factor of four, with the rate increasing by approximately a factor of four as well. $[S_2O_8^{2-}]$ also doubles by examining reaction 5 to 4. Thus when $[I^-]$ is constant, the reaction is first-order in $[S_2O_8^{2-}]$, with $m = 1$. ✓

3. Write a rate law for this reaction that is consistent with your conclusions.

$rate = k[S_2O_8^{2-}][I^-]$ ✓

4. Calculate the rate constant k for each of the five reactions in Part A. Average the five values.

Rxn.	k
1	3.59×10^{-3}
2	3.32×10^{-3}
3	3.21×10^{-3}
4	2.11×10^{-3}
5	3.70×10^{-3}
Units →	$L\cdot mol^{-1}s^{-1}$

$$k = \frac{rate}{[S_2O_8^{2-}][I^-]}$$

$$k_1 = \frac{5.75\times10^{-6}\ mol\cdot L^{-1}s^{-1}}{(4.00\times10^{-2}\ mol/L)(4.00\times10^{2}\ mol/L)} = 3.59\times10^{-3}\ L\cdot mol^{-1}s^{-1}$$

$$k_2 = \frac{1.06\times10^{-5}\ mol\cdot L^{-1}\cdot s^{-1}}{(8.00\times10^{-2}\ mol/L)(4.00\times10^{-2}\ mol/L)} = 3.32\times10^{-3}\ L\cdot mol^{-1}s^{-1}$$

$$k_3 = \frac{2.56\times10^{-6}\ mol\cdot L^{-1}s^{-1}}{(2.00\times10^{-2}\ mol/L)(4.00\times10^{-2}\ mol/L)} = 3.21\times10^{-3}\ L\cdot mol^{-1}s^{-1}$$

$$k_4 = \frac{3.38\times10^{-6}\ mol\cdot L^{-1}s^{-1}}{(8.00\times10^{-2}\ mol/L)(2.00\times10^{-2}\ mol/L)} = 2.11\times10^{-3}\ L\cdot mol^{-1}s^{-1}$$

$$k_5 = \frac{2.96\times10^{-6}\ mol\cdot L^{-1}s^{-1}}{(8.00\times10^{-2}\ mol/L)(1.00\times10^{-2}\ mol/L)} = 3.70\times10^{-3}\ L\cdot mol^{-1}s^{-1}$$

✓ range is too large

−1

Average Temperature 22.0°C

(This is the average of the five average temperatures.)

Average k $3.19\times10^{-3}\ L\cdot mol^{-1}s^{-1}$

a bit low

−1

Part B. Dependence of Reaction Rate on Catalyst

Catalyst Unknown No. 501

	Time Seconds	Temp. °C	Initial [I^-] (mol L^{-1})	Initial [$S_2O_8^{2-}$] (mol L^{-1})	Rate (mol L^{-1} s^{-1})	k ($L\cdot mol^{-1}s^{-1}$)
6(a)	60 ~~59.6~~ ~~59.6~~	21.9 21.9 21.9	4.00×10^{-2}	4.00×10^{-2}	8.33×10^{-6}	5.21×10^{-3}
6(b)	57 ~~56.5~~	22.0 21.8 21.9	4.00×10^{-2}	4.00×10^{-2}	8.77×10^{-6}	5.48×10^{-3}
6(c)	57 ~~56.7~~	21.9 21.9 21.9	4.00×10^{-2}	4.00×10^{-2}	8.77×10^{-6}	5.48×10^{-3}

Average Time 58 s Average Temperature 21.9 °C

Average k 5.39×10^{-3} Units $L\cdot mol^{-1}s^{-1}$

$$\text{Initial }[I^-] = \frac{(0.200\text{ mol/L})(10.0\times10^{-3}L)}{50.0\times10^{-3}L} = 4.00\times10^{-2}\text{ mol/L}$$

$$\text{Initial }[S_2O_8^{2-}] = \frac{(0.100\text{ mol/L})(20.0\times10^{-3}L)}{50.0\times10^{-3}L} = 4.00\times10^{-2}\text{ mol/L}$$

reaction (a):

$$\text{rate} = \frac{\Delta[S_2O_8^{2-}]}{\Delta t} = \frac{5.00\times10^{-4}\text{ mol/L}}{60\text{ s}} = 8.33\times10^{-6}\text{ mol}\cdot L^{-1}s^{-1}$$

reaction (b), (c):

$$\text{rate} = \frac{\Delta[S_2O_8^{2-}]}{\Delta t} = \frac{5.00\times10^{-4}\text{ mol/L}}{57\text{ s}} = 8.77\times10^{-6}\text{ mol}\cdot L^{-1}s^{-1}$$

$$\text{rate} = k[S_2O_8^{2-}][I^-]$$

$$k_a = \frac{8.33\times10^{-6}\text{mol}\cdot L^{-1}\cdot s^{-1}}{(4.00\times10^{-2}\text{mol/L})(4.00\times10^{-2}\text{mol/L})} = 5.21\times10^{-3}\ L\cdot mol^{-1}s^{-1}$$

$$k_b = \frac{8.77\times10^{-6}\text{mol}\cdot L^{-1}s^{-1}}{(4.00\times10^{-2}\text{mol/L})(4.00\times10^{-2}\text{mol/L})} = 5.48\times10^{-3}\ L\cdot mol^{-1}s^{-1}$$

$$k_c = 5.48\times10^{-3}\ L\cdot mol^{-1}s^{-1}$$

RP

Part C. Effect of Temperature on Reaction Rate

Approx. Temperature, °C	0		Room		40	
Measured Temperature, °C	6.0	3.6	22.0	22.2	35.0	35.9
Average	~~[illegible]~~ 4.8		22.1		35.5	
Time	~~417.9~~s 418s		87s		~~33.5~~s 34s	
Rate, mol L^{-1} s^{-1}	1.20×10^{-6} mol·$L^{-1}s^{-1}$		5.75×10^{-6} mol·$L^{-1}s^{-1}$		1.47×10^{-5} mol·$L^{-1}s^{-1}$	
Relative Rate x 100	100		480		1229	
log of (Relative Rate × 100)	2.00		2.68		3.09	
Temperature, K	278 K		295 K		309 K	
1 / T, K^{-1}	3.60×10^{-3} K^{-1}		3.39×10^{-3} K^{-1}		3.24×10^{-3} K^{-1}	

Draw the graph. Calculate the slope. Calculate E_a

$$\text{rate}_{0°C} = \frac{\Delta[S_2O_8^{2-}]}{\Delta t} = \frac{5.00 \times 10^{-4}\ \text{mol/L}}{418\ \text{s}} = 1.20 \times 10^{-6}\ \text{mol} \cdot L^{-1} s^{-1}$$

$$\text{rate}_{\text{room}} = \frac{5.00 \times 10^{-4}\ \text{mol/L}}{87\ \text{s}} = 5.75 \times 10^{-6}\ \text{mol} \cdot L^{-1} s^{-1}$$

$$\text{rate}_{40°C} = \frac{5.00 \times 10^{-4}\ \text{mol/L}}{34\ \text{s}} = 1.47 \times 10^{-5}\ \text{mol} \cdot L^{-1} s^{-1}$$

$$\text{rel.rate}_{0°C} = \frac{\text{rate}_{0°C}}{\text{rate}_{0°C}} \times 100 = 100$$

$$\text{rel.rate}_{\text{room}} = \frac{\text{rate}_{\text{room}}}{\text{rate}_{0°C}} \times 100 = \frac{5.75 \times 10^{-6}}{1.20 \times 10^{-6}} \times 100 = 480$$

$$\text{rel.rate}_{40°C} = \frac{\text{rate}_{40°C}}{\text{rate}_{0°C}} \times 100 = \frac{1.47 \times 10^{-5}}{1.20 \times 10^{-6}} \times 100 = 1229$$

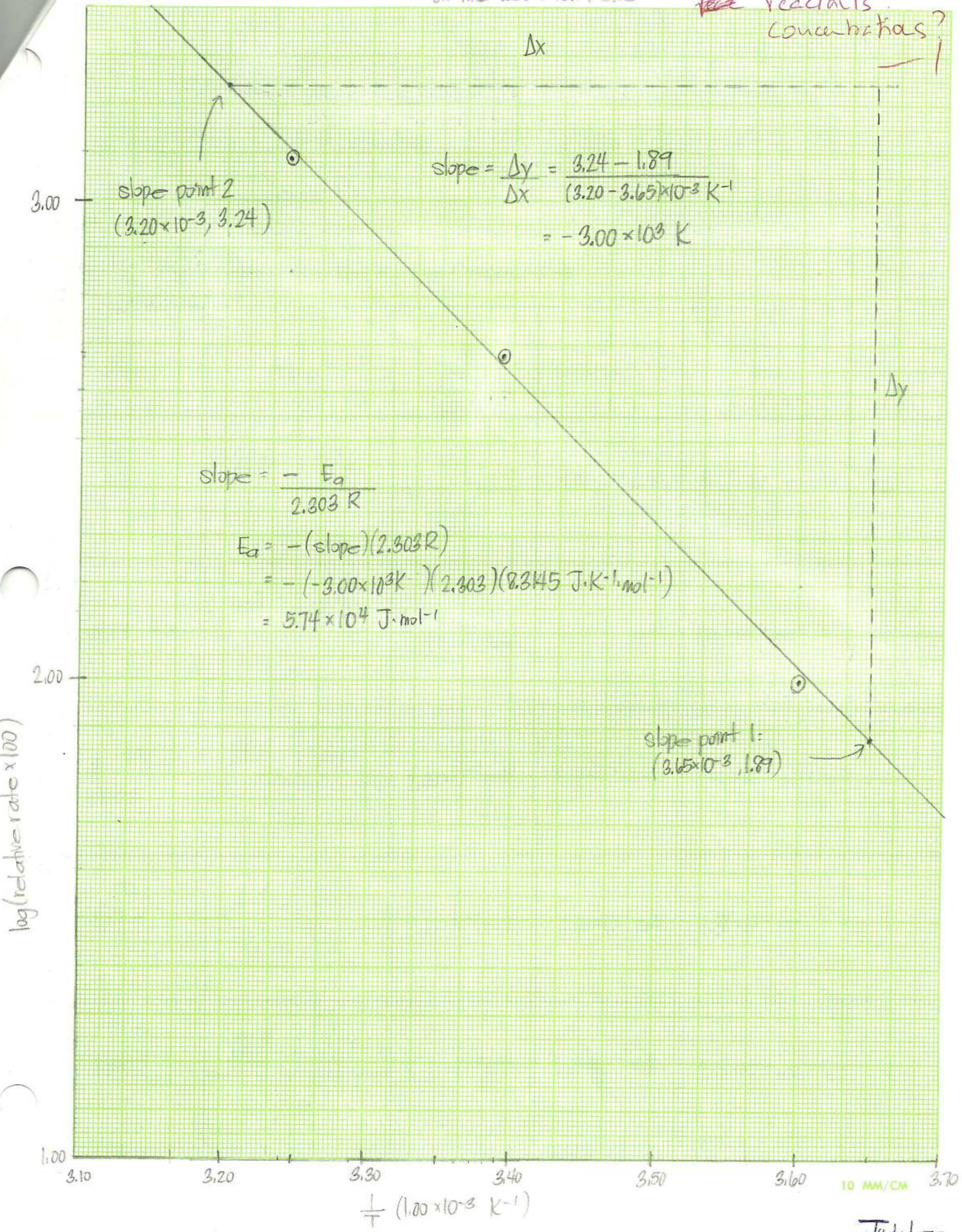
(slope + Ea calculations on next page)
on graph

Slope = -3.00×10^{3} K ✓

E_a = 5.74×10^{4} J $\cdot$ mol^{-1} ✓

Determining the Activation Energy, E_a from the Effects of Temperature at 4.8°C, 22.1°C, 35.5°C on the Reaction Rate

reactants? concentrations?



$$\text{slope} = \frac{\Delta y}{\Delta x} = \frac{3.24 - 1.89}{(3.20 - 3.65)\times10^{-3}\ K^{-1}}$$
$$= -3.00\times10^{3}\ K$$

$$\text{slope} = \frac{-\ E_a}{2.303\ R}$$

$$E_a = -(\text{slope})(2.303 R)$$
$$= -(-3.00\times10^{3} K)(2.303)(8.3145\ J\cdot K^{-1}\cdot mol^{-1})$$
$$= 5.74\times10^{4}\ J\cdot mol^{-1}$$

Judy Lee