

---

## Investigation 11: Measuring Enthalpy of Reactions

### Aim:

The objective of this investigation is to measure the heat evolved or absorbed in some physical and chemical changes.

### Introduction:

This investigation is designed to determine the enthalpy change for the displacement reaction between iron and aqueous copper sulphate. The equation of the reaction is  $\text{CuSO}_{4(aq)} + \text{Zn}_{(s)} \rightarrow \text{ZnSO}_{4(aq)} + \text{Cu}_{(s)}$ . By placing both zinc and aqueous copper sulphate in the calorimeter and having the reaction begin, the temperature change could be observed and recorded by the thermometer. Hence, heat energy could be calculated through the equation  $\text{Heat energy (kJ)} = C \cdot m \cdot \Delta T$  where  $C$  is specific heat capacity,  $m$  is the mass, and  $\Delta T$  is the temperature change while the specific heat capacity of water and dilute aqueous solutions as  $4.183 \text{ kJ kg}^{-1} \text{K}^{-1}$ . Then, the enthalpy change ( $\text{kJ} \cdot \text{mol}^{-1}$ ) could be determined as well by having heat energy be divided by the mole of the limited reagent.

### Apparatus:

1 stirring rod	1M aqueous copper(II) sulphate
Powder zinc	Electronic balance ( $\Delta = \pm 0.001\text{g}$ )
2 cylinders (50ml, 2ml, $\Delta = \pm 1\text{ml}$ )	Stop watch
Thermometer ( $100.0^\circ\text{C}$ , $0.2^\circ\text{C}$ , $\Delta = \pm 0.1^\circ\text{C}$ )	Polystyrene (foam) cup
1 beaker (100ml)	3 small beakers (50ml)

### Method:

1. Clean calorimeter, cylinders and beakers. Make sure they are dry before use them.
2. Take a calorimeter and use a cylinder to put  $30\text{ml} \pm 1\text{ml}$  of 1 M aqueous copper (II) sulphate in it.
3. Weigh out accurately about 4g of powdered zinc into a small beaker.
4. Stir and record the temperature of copper sulphate at half-minute intervals for 2 minutes and then add the powdered zinc into the calorimeter.
5. Record the temperature at half-minute intervals until the temperature has been falling for ten consecutive readings.
6. Repeat steps 1-5 for another two times.

### Raw Data collection:

### Observation

1. When the reaction began, water vapor formed on the plastic cover of the calorimeter, heat could be felt on the cover, and temperature indicated by the thermometer rose.

2. Copper sulfate solution was blue and zinc powder was grey.
3. The blue color of copper sulfate solution became colorless in the experiment.

**Table 1: Mass of zinc**

1 <sup>st</sup> trial Mass of zinc ( $\Delta = \pm 0.001\text{g}$ )	2 <sup>nd</sup> trial Mass of zinc ( $\Delta = \pm 0.001\text{g}$ )	3 <sup>rd</sup> trial Mass of zinc ( $\Delta = \pm 0.001\text{g}$ )
3.998	4.000	4.002

**Table 2: Temperature recorded at each time interval**

Time(s) ( $\Delta = \pm 1\text{s}$ )	1 <sup>st</sup> trial Temperature ( $\Delta = \pm 0.1^\circ\text{C}$ )	2 <sup>nd</sup> trial Temperature ( $\Delta = \pm 0.1^\circ\text{C}$ )	3 <sup>rd</sup> trial Temperature ( $\Delta = \pm 0.1^\circ\text{C}$ )
0	24.0	23.5	24.0
30	29.8	42.9	52.0
60	31.1	49.2	63.9
90	41.0	52.1	63.0
120	42.1	53.9	61.0
150	42.3	55.1	60.0
180	43.9	56.0	59.1
210	44.5	56.1	58.5
240	50.1	55.6	57.9
270	50.5	54.8	57.0
300	51.0	54.2	56.2
330	52.1	53.5	55.2
260	52.7	53.1	54.9
390	52.4	53.0	54.6
420	52.1	52.3	53.7
450	52.0	52.0	53.0
480	52.0	51.5	52.5
510	52.0	51.0	51.9
540	52.0	50.8	51.3
570	52.0	50.8	51.0
600	52.0	50.8	50.5
630	52.0	50.8	50.5

**Table 3: Other related data**

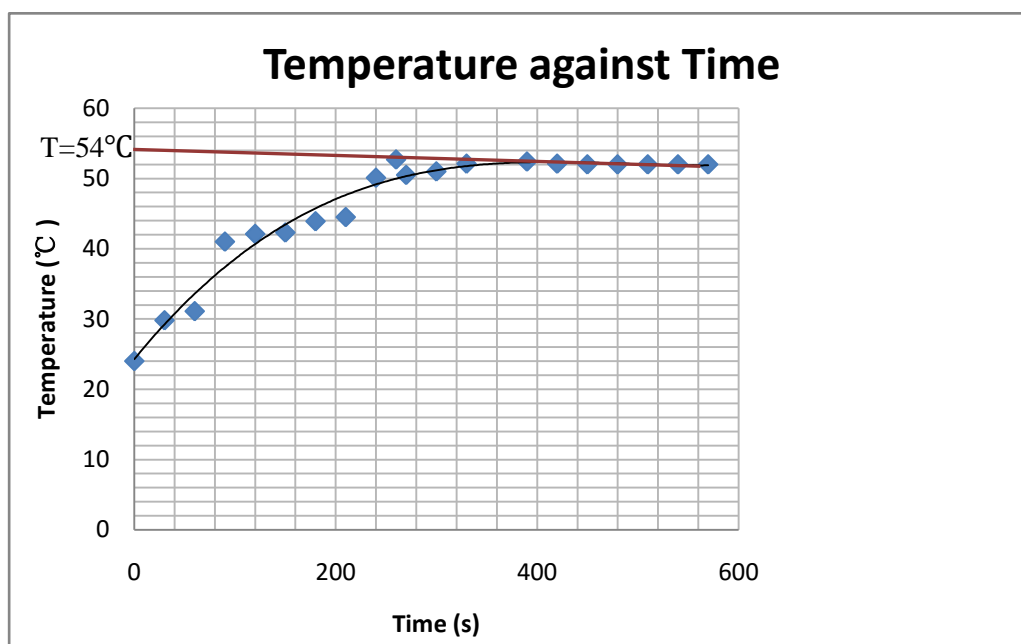
Volume of $\text{CuSO}_4$	30ml $\pm$ 1ml
Concentration of $\text{CuSO}_4$	1mol/L

### Calculation:

Sample calculation of Trial 1:

Determine highest temperature:

Based on the temperature recorded at each time interval for trial 1, a graph of temperature against time is shown below and best straight line going through the cooling section (all points from the highest temperature to the final temperature) intersects with the temperature axis at  $54.0^{\circ}\text{C}$ . Here, the errors are too small to show on the graph. Hence, the highest temperature that would have occurred if the reaction had been instantaneous and heat had not been lost to the environment is  $54.0 \pm 0.1^{\circ}\text{C}$



Calculate temperature change:

$$T_{\text{change}} = T_{\text{highest}} - T_{\text{initial}} = 54.0 - 24.0 = 30.0^{\circ}\text{C}$$

$$\Delta T_{\text{change}} = \pm(0.1 + 0.1) = \pm 0.2^{\circ}\text{C}$$

$$\therefore T_{\text{change}} = 30.0 \pm 0.2^{\circ}\text{C}$$

Calculate heat energy:

(assume the density of dilute copper (II) sulphate is the same as the density of water which is  $1\text{g} \cdot \text{cm}^{-3}$  and the heat capacity of dilute copper (II) sulphate is the same as that of water which is  $4.183\text{kJ} \cdot \text{kg}^{-1} \cdot ^{\circ}\text{C}^{-1}$ .)

$$\text{mass of } \text{CuSO}_{4(aq)} = 30\text{cm}^3 \times 1\text{g} \cdot \text{cm}^{-3} = 30\text{g}$$

$$\Delta \text{mass} = \pm \left( \frac{1}{30} + \frac{0}{1} \right) \times 30 \approx \pm 1\text{g}$$

$$\text{Heat energy} = C \cdot \text{mass of } \text{CuSO}_{4(aq)} \cdot \Delta T$$

$$\text{Heat energy} = 30 \times 10^{-3}\text{kg} \times 4.183\text{kJ} \cdot \text{kg}^{-1} \cdot ^{\circ}\text{C}^{-1} \times 30.0^{\circ}\text{C} \approx 3.8\text{kJ}$$

$$\Delta \text{Heat energy} = \pm \left( \frac{1 \times 10^{-3}}{30 \times 10^{-3}} + \frac{0}{4.183} + \frac{0.2}{30.0} \right) \times 3.8 \approx \pm 0.2 \text{ kJ}$$

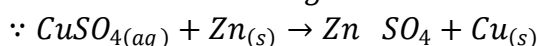
- ∴ Temperature rises through the experiment and hence this reaction is exothermic
- ∴ Heat energy is given out and its value has to be negative
- ∴ Heat energy =  $-3.8 \pm 0.2 \text{ kJ}$

Calculate limiting reagent:

$$n_{\text{CuSO}_4} = 1 \text{ mol} \cdot \text{l}^{-1} \times 30 \times 10^{-3} \text{ l} = 3 \times 10^{-2} \text{ mol}$$

$$\Delta n_{\text{CuSO}_4} = \pm \left( \frac{0}{1} + \frac{1 \times 10^{-3}}{30 \times 10^{-3}} \right) \times 3 \times 10^{-2} = \pm 1 \times 10^{-3} \text{ mol}$$

$$n_{\text{Zinc}} = \frac{\text{mass}}{\text{molar mass}} = \frac{3.998 \text{ g}}{65.37 \text{ g} \cdot \text{mol}^{-1}} \approx 0.06116 \text{ mol}$$



$$n_{\text{CuSO}_4} : n_{\text{Zinc}} = 1 : 1 \text{ (should be)}$$

$$\text{While the real ratio } n_{\text{CuSO}_4} : n_{\text{Zinc}} = 0.03 : 0.06116 < 1 : 1$$

∴ *CuSO<sub>4</sub> is limiting reagent.*

Calculate enthalpy change:

$$\text{Enthalpy change} = \frac{\text{Heat energy}}{\text{mole of limiting reagent}} = \frac{-3.8 \text{ kJ}}{3 \times 10^{-2} \text{ mol}} \approx -1 \times 10^2 \text{ kJ} \cdot \text{mol}^{-1}$$

$$\Delta \text{Enthalpy change} = \pm \left( \frac{0.2}{-3.8} + \frac{1 \times 10^{-3}}{3 \times 10^{-2}} \right) \times (-1 \times 10^2) \approx \pm 0.9 \text{ kJ} \cdot \text{mol}^{-1}$$

$$\therefore \text{Enthalpy change} = -1 \times 10^2 \pm 9 \times 10^{-1} \text{ kJ} \cdot \text{mol}^{-1}$$

By using the same method to calculate the temperature change, mole of limiting reagent, and heat energy, enthalpy changes of the other two trials are shown in the table below.

**Table 4: Calculation results and enthalpy changes of 3 trials**

	Highest T (°C)	T change (°C)	Heat energy (kJ)	Mole of limiting reagent (mol)	Enthalpy change $\text{kJ} \cdot \text{mol}^{-1}$
Trial 1	$54.0 \pm 0.1$	$30.0 \pm 0.2$	$3.8 \pm 0.2$	$3 \times 10^{-2} \pm 1 \times 10^{-3}$	$-1 \times 10^2 \pm 9 \times 10^{-1}$
Trial 2	$58.9 \pm 0.1$	$34.9 \pm 0.2$	$4.2 \pm 0.2$	$3 \times 10^{-2} \pm 1 \times 10^{-3}$	$-1 \times 10^2 \pm 5$
Trial 3	$66.0 \pm 0.1$	$42.0 \pm 0.2$	$5.3 \pm 0.2$	$3 \times 10^{-2} \pm 1 \times 10^{-3}$	$-2 \times 10^2 \pm 1 \times 10^1$

Then, the mean value of the enthalpy changes of 3 trials could be calculated: (during the calculation, original data are used instead of the estimated value of these original data)

$$\begin{aligned}\text{mean enthalpy change} &= \frac{(-125.5) + (-145.99) + (-175.69)}{3} = \frac{-447.18}{3} \\ &= -149.06 \approx -1 \times 10^2 \text{ kJ} \cdot \text{mol}^{-1} \\ \Delta \text{ mean} &= \pm \left( \frac{9 \times 10^{-1} + 5 + 10}{-447.18} + \frac{0}{3} \right) \times -1 \times 10^2 \approx \pm 4 \text{ kJ} \cdot \text{mol}^{-1} \\ \therefore \text{Enthalpy change of this reaction is } &-1 \times 10^2 \pm 4 \text{ kJ} \cdot \text{mol}^{-1}\end{aligned}$$

### Conclusion:

The enthalpy change of the reaction between copper (II) sulphate and zinc is  $-1 \times 10^2 \pm 4 \text{ kJ} \cdot \text{mol}^{-1}$  which is got from the experiment while the theoretic value is  $-218 \text{ kJ} \cdot \text{mol}^{-1}$  from “General Chemistry-Principle & Modern Applications”, Sixth Edition, Petrucci and Harwood.

$$\text{Percentage discrepancy} = \left| \frac{(-1 \times 10^2) - (-218)}{(-218)} \right| \times 100\% \approx 54.1\%$$

The discrepancy between the experimental value and the theoretical value indicates there must be many systematic errors existed during the investigation.

Meanwhile, the percentage error of the experimental result is

$$\Delta\% = \frac{4}{100} = 4\%$$

As this percentage error is much smaller than the percentage discrepancy, this circumstance means that random error alone cannot account for the discrepancy and hence some systematic errors must have occurred. These errors need to be improved, such as the heat absorbed by the calorimeter and heat exchange with the surroundings, and they would be discussed in the next part.

Furthermore, a chemistry explanation could be provided to explain the exothermic feature of this reaction between zinc and copper (II) sulfate. Regarding to the nature of chemical reactions, as zinc reacted with copper (II) sulfate, bond between  $\text{Cu}^{2+}$  and  $\text{SO}_4^{2-}$  was broken and that between  $\text{Zn}^{2+}$  and  $\text{SO}_4^{2-}$  was formed, due to the fact that zinc is more reactive than copper (II). The process of bond broken is endothermic, absorbing energy, while the process of bond formation is exothermic, releasing energy. Consequently, as the amount of energy released is more than that of energy absorbed, this reaction is exothermic.

---

**Evaluation and Improvement:****Table 5:**

Errors cause uncertainties	Improvement
The temperature recorded at each 30 seconds interval may not be accurate because there is reaction time for a person to record the temperature after that person's partner calls the time.	This uncertainty is acceptable because everyone has reaction time. An improvement could be repeating the experiment for more than 3 times to minimize the uncertainty caused by a person's reaction time.
There were several intervals of sudden increase or sudden decrease in temperature from the 3 groups of data. This uncertainty was caused by the fluctuation of the speed of stirring the solution. A sudden increase in temperature means a sudden increase in the speed of stirring the solution.	This uncertainty could be minimized by controlling a person's speed of stirring the solution carefully and slowly since a person could control a slower speed of stirring better. Meanwhile, repeating the experiment for more than 3 times could also minimize this error.
There must be some heat exchanged with the surrounding or heat absorbed by the calorimeter since the material of the calorimeter in the school laboratory is Polystyrene (foam) cup which may not have the best heat insulating ability. Consequently, the experimental value of the absolute value of the heat energy may be smaller than that of theoretical value.	Since the material could not be changed, more layers of Polystyrene cup could be used in the calorimeter and hence have a better heat insulating ability and better accuracy as well.
Since we had to assume that the solution of copper (II) sulphate given to us had the exact concentration stated and had the heat capacity same as the heat capacity of water. However, since it was prepared by a human, there is no doubt there would be a margin of error. Also, as the density of copper (II) sulfate was assumed to be the same as that of water, the uncertainty of the calculation must be greater than that we calculated as well. Hence, the experimental result may be more inaccurate.	In order to ensure the uncertainty of the concentration of copper sulphate used in the experiment, it would be better to make 1 mole per liter copper sulphate by the students. As a result, its uncertainty is the addition of the uncertainties of the cylinders used to measure the volumes of water and copper sulphate. Meanwhile, it can also be improved by researching the real heat capacity of copper (II) sulfate and the density of it in the text books.
Since there was much water vapor formed on the plastic cover of the calorimeter, this phenomenon indicated that loss of energy occurs due to the evaporation of the water in the	In order to minimize the energy loss created by the evaporation of water, a higher pressure could be applied to the calorimeter. Hence, the experiment result could be more accurate and approaching

---

experiment. As a result, the absolute value of the enthalpy change measured in this experiment was smaller than that of the theoretical value.	to the theoretical value.
--	---------------------------