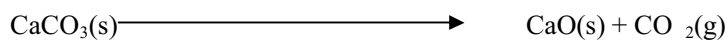


Chemistry Coursework – Assessed Practical

Introduction

Calcium Carbonate, CaCO_3 decomposes with heat, to produce Calcium Oxide, CaO and Carbon Dioxide, CO_2 .



The object of this practical exercise is to determine the enthalpy change for this reaction by an indirect method based on Hess' Law.

Results Table

Experiment 1:

Mass of CaCO_3 and weighing bottle	3.47g
Mass of empty weighing bottle	1.08g
Mass of CaCO_3 used	2.39g
Temperature of acid initially	22°C
Temperature of solution after mixing	24°C
Temperature change during reaction	2°C

Experiment 2:

Mass of CaO and weighing bottle	2.58g
Mass of empty weighing bottle	1.00g
Mass of CaO used	1.58g
Temperature of acid initially	22°C
Temperature of solution after mixing	27°C
Temperature change during reaction	5°C

Calculating the Enthalpy Change

Experiment 1: Calculating the ΔH_1 for the reaction between CaCO_3 and HCl

Using the formula:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

We can work out the mass of liquid (HCl)

$$\begin{aligned}\text{Mass} &= \text{Density} \times \text{volume} \\ &= 2.0 \times 0.05 \\ &= 0.1 \text{ grams}\end{aligned}$$

Using the calculated mass, we can work out the total energy transfer using the formula:

$$\begin{aligned}\text{Energy Transfer} &= \text{Mass of Liquid} \times \text{specific heat capacity} \times \text{change in temperature} \\ &= 0.1 \times 4.2 \times (24-22) \\ &= -0.84 \text{ J}\end{aligned}$$

From the results table, we can see that the temperature increased from 22°C to 24°C, so we can deduce that the reaction was EXOTHERMIC.

To work out ΔH_1 we need to find the energy transfer per 1 mole of calcium carbonate, but first I need to calculate the number of moles in calcium carbonate:

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molar Mass}}$$

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$$\begin{aligned} & \frac{\text{Mr}}{= 2.39} \\ & \frac{100.1}{= 0.02387612388 \text{ moles}} \\ & = 0.024 \text{ moles (3 d.p.)} \end{aligned}$$

Therefore,

$$\begin{aligned} \Delta H_1 &= \frac{-0.84}{0.024} \\ &= -35 \text{ J mol}^{-1} \end{aligned}$$

Experiment 2: Calculating the ΔH_2 for the reaction between CaO and HCl.

Using the formula:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

We can work out the mass of liquid (HCl)

$$\begin{aligned} \text{Mass} &= \text{Density} \times \text{volume} \\ &= 2.0 \times 0.05 \\ &= 0.1 \text{ grams} \end{aligned}$$

We can now put the calculated mass into the energy transfer equation, as below:

$$\begin{aligned} \text{Energy Transfer} &= \text{Mass of Liquid} \times \text{specific heat capacity} \times \text{change in temperature} \\ &= 0.1 \times 4.2 \times (27-22) \\ &= -2.1 \text{ J} \end{aligned}$$

The energy transfer has a negative value, as the reaction was exothermic. This can be seen by the increase in temperature after the reaction.

To work out ΔH_1 we need to find the energy transfer per 1 mole of calcium oxide, but first I need to calculate the number of moles in calcium oxide:

$$\begin{aligned} \text{Number of moles} &= \frac{\text{Mass}}{\text{Mr}} \\ &= \frac{1.58}{56.1} \\ &= 0.02816399287 \text{ moles} \\ &= 0.028 \text{ moles (3 d.p.)} \end{aligned}$$

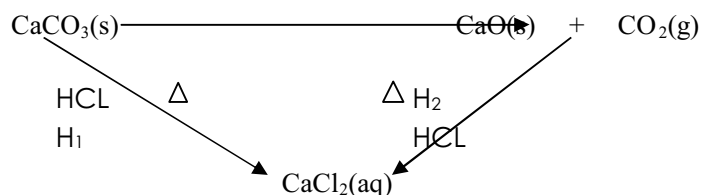
Therefore,

$$\begin{aligned} \Delta H_1 &= \frac{-2.1}{0.028} \\ &= -75 \text{ J mol}^{-1} \end{aligned}$$

Hess' Law states that 'for a given overall chemical reaction, the overall energy is the same whether the reaction takes place in a single step or via a series of steps' – provided that all the measurements are made under the same conditions. This means we can use this theory to work out the enthalpy change of the thermal decomposition of CaCO_3 .

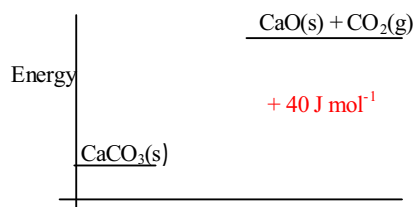
$$\Delta H_3$$

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$$\begin{aligned}\Delta H_3 &= \Delta H_1 - \Delta H_2 \\ &= -35 - (-75) \\ &= 40 \text{ J mol}^{-1}\end{aligned}$$

I can draw an energy profile to see the changes more clearly:



The energy profile shows the reaction between the products and reactants was *endothermic*. This means that the total enthalpy – the energy content- of the system goes up.

Evaluation

The analysis of the results shows that the calculations produced reliable answers, while producing integers. However, after looking over my method, I can see a few minor errors that could have altered the results given.

In the experiment, one or two minor misjudgements could have been made; although such mistakes will not have been sufficient enough to change the results. The top pan balance, used for weighing out the calcium carbonate and calcium oxide, is a very accurate method of measuring quantities. However, top pan balances only measure to two decimal places. On a balance like this the second decimal place is not reliable, giving an error of ± 0.01 g.

CaCO₃ $(0.01/2.39) \times 100 = 0.42\%$ chance of error

CaO $(0.01/ 1.58) \times 100 = 0.63\%$ chance of error.

To overcome this problem, it is important to weigh out the solids to the required weigh; take them off the scale; place them back on the scale and check the weight. If the two weights are very different from each other, it may be best to weigh the solids again. However the calculated chances of errors are not high enough to cause any concern for the accuracy of my results.

One of the major concerns about the accuracy of my results was the ability of the thermometer to read the temperature to the required level. The thermometer ranged from 0°C to 100°C, in graduations to 1°C, so I could read the thermometer accurately and round the temperatures to the nearest centigrade. In the thermometer, there is a reservoir which is where the thermometer detects the heat. The liquid in the reservoir expands as the temperature of the liquid increases. The expansion of the liquid corresponds to the temperature of the liquid; i.e. the more the liquid expands, the greater the temperature. However, the liquid reservoir (red bulb) in the thermometer was not completely submerged in the liquid in the beaker; causing the recorded temperature to not be 100% accurate. This will cause the liquid to expand at a slower rate, and the expansion will be less; resulting in a very slightly lower temperature. To try and overcome this problem, I could use a thinner, narrower beaker where the liquid level will be much higher. The beaker I used was an ordinary glass beaker, without a lid. Insulation of the beaker was limited, which led to a greater percentage of the heat given off, was

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lost to its surroundings. To make the experiment more accurate, there is the possibility of using a more insulated will reduce the percentage of energy lost to the surroundings. There is another possibility of using a beaker with a lid. Heat will be lost at a faster rate, which results in a higher temperature after the reaction. The speed of reaction will also affect the amount of heat lost to its surroundings. This is important in this experiment, as experiment 1 had a fairly fast rate of reaction, while experiment 2 had a very slow rate of reaction. The slower reaction may cause more heat to be lost during the course of the reaction. To make the experiment more accurate I could make the solids into a more powdered form, which will speed up the reaction as a result of an increase in surface area. This means there is a lesser chance of energy being lost.

All these 3 errors are significant to the experiment, as they all can contribute to an increase/ decrease of the temperature of the mixture. This, in turn, will distort the calculated enthalpy changes, and they will be incorrect. It is known that beakers can be not be very accurate, as the measurements are in 50cm^3 graduations; I decided to use a measuring cylinder to measure the acid, which range from 0cm^3 to 250cm^3 , with 1.0cm^3 graduations. Therefore the measurements of volume will be rounded to the nearest 1.0cm^3 , giving a degree of inaccuracy. When measuring the acid out, I need to make sure the meniscus is at the bottom of the measurement line.

All these four errors are significant to the experiment, as they can all contribute to an increase or decrease of the temperature of the experiment. The most significant error in the method was the fact that the temperature is not correctly identified by the thermometer. I believe this to be the most important factor, as this is the main way of detecting the temperature. If the temperature is slightly out, the end result will be slightly out.

Overall I think that experiment 1; involving CaCO_3 was the more reliable of the two experiments. The main reason behind this decision was the fact that the reaction is very quick, leaving less time for any heat to be lost to the environment and its surroundings.