

Results

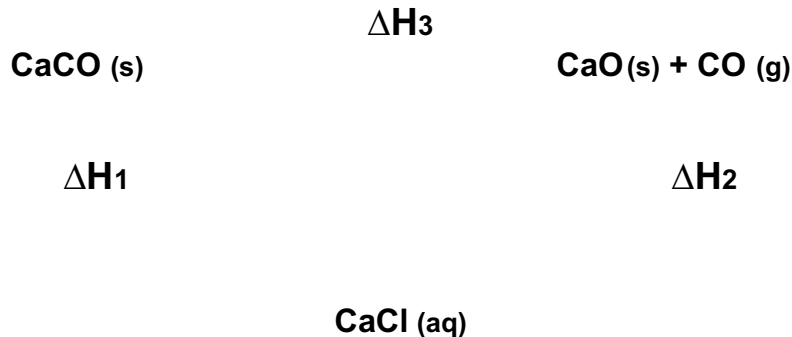
CaCO	Test 1	Test 2	Average
Mass of CaCO Used :-	2.48g	2.49g	2.485g
Temperature of HCl Initially :-	20°C	20°C	20°C
Temperature of HCl after mixing with CaCO: -	22°C	22°C	22°C
Total Temperature Change :-	2°C	2°C	2°C

CaO	Test 1	Test 2	Average
Mass of CaO Used :-	1.40g	1.40g	1.40g
Temperature of HCl Initially :-	19°C	20°C	19.5°C
Temperature of HCl after mixing with CaO: -	38°C	36°C	37°C
Total Temperature Change :-	18°C	16°C	17°C

In order to calculate the enthalpy change of Calcium Carbonate to Calcium Oxide, which is exceedingly hard to control and measure the energy change we must use Hess' law which states: -

"The total enthalpy change for a chemical reaction is independent of the route by which the reaction takes place, provided initial and final conditions are the same"

So therefore on adding Hydrochloric acid, which reacts readily with both, Calcium Carbonate and Calcium Oxide they both form Calcium Chloride creating an enthalpy energy triangle.



Therefore:-

$$\Delta H_1 - \Delta H_2 = \Delta H_3$$

But we cannot enter our results directly into Hess' equation, we must convert them into Enthalpy energies. With the enthalpy energy values for Calcium Carbonate and Calcium Oxide only then can we calculate the enthalpy change for Calcium Carbonate and Calcium Oxide.

Enthalpy Calculations

There are a few equations I must use to obtain my final answer firstly I must calculate the energy transferred in the reaction using the following equation: -

Equation 1

Mass of liquid(g) x specific heat capacity of
aqueous solution(Jg K) x temperature rise(°C) = Energy Transferred (J)

OR

$$\Delta H = -M \times C_p \times \Delta T$$

But this doesn't calculate the energy transferred per mole, to do this I must use:-

Equation 2

$$\frac{\text{Mass of Chemical used(g)}}{\text{Molecular weight of Molecule}} = \text{No. of Moles (mol)}$$

Lastly we must convert the energy transferred into moles to do this: -

Equation 3

$$\text{Enthalpy Energy} \quad \frac{\Delta H}{\text{No of Moles}_-} \times \quad =$$

CaCO₃ - ΔH1

1. Enter results into equation 1

$$\Delta H = -M \times C_p \times \Delta T$$

$$\rightarrow \Delta H = -50 \times 4.2 \times 2 = -420\text{J}$$

2. Enter results into equation 2

$$\frac{\text{Mass of Chemical used(g)}}{\text{Molecular weight of Molecule}} = \text{No. of Moles (mol)}$$

$$\rightarrow \frac{2.485}{\underset{\text{Ca}}{40.1} + \underset{\text{C}}{12} + \underset{\text{O}}{(16 \times 3)}} = 0.024825174 \text{ mol}$$

3. Enter results into equation 3

$$\Delta H \times \frac{1}{\text{No of Moles}_-} \times = \text{Enthalpy Energy (Jmol)}$$

$$\rightarrow -420\text{J} \times \frac{1}{0.024825174} = \underline{\underline{-16918.31042 \text{ Jmol}}}$$

CaO - ΔH2

1. Enter results into equation 1

$$\Delta H = -M \times C_p \times \Delta T$$

$$\rightarrow \Delta H = -50 \times 4.2 \times 17 = -3570\text{J}$$

2. Enter results into equation 2

$$\frac{\text{Mass of Chemical used(g)}}{\text{Molecular weight of Molecule}} = \text{No. of Moles (mol)}$$

$$\rightarrow \frac{1.4}{40.1 + 16} = 0.024955436 \text{ mol}$$

3. Enter results into equation 3

$$\Delta H \times \frac{1}{\text{No of Moles}_-} = \text{Enthalpy Energy (Jmol)}$$

$$\rightarrow -3570\text{J} \times \frac{1}{0.024955436} = \underline{\underline{-143055.0041 \text{ Jmol}}}$$

Now with my two results (enthalpy energies) for Calcium Carbonate and Calcium Oxide I can now work out the missing Enthalpy change using Hess's law:-

$$\Delta H_1 - \Delta H_2 = \Delta H_3$$

$$\rightarrow -16918.31042 - (-143055.0041) = \underline{\underline{126136.6937 \text{ Jmol}}} \\ = \underline{\underline{126 \text{ kJmol to 3sf}}}$$

With my result and the result produced by the data book I can calculate my percentage error. The information produced in the data book is under Standard Conditions. These conditions are: -

- A pressure of 100 kilopascals
- A temperature of 298K
- The reactants must be in their natural physical state
- All solutions must have a concentration of 1 mol dm⁻³

With these accurate results from the data book I can calculate my percentage error:-

Data book Information on
"The Standard Enthalpy Change of Formation"

$$\Delta H_1 - \text{Calcium Carbonate} = -1207 \text{ kJmol}$$

$$\Delta H_2 - \text{Calcium Oxide} = -635 \text{ kJmol}$$

$$\Delta H_2 - \text{Carbon Dioxide} = -394 \text{ kJmol}$$

So on putting our Standard Enthalpy results into Hess's equation we get a Standard Enthalpy Change for Calcium Carbonate to Calcium Oxide: -

$$\Delta H_1 - \Delta H_2 = \Delta H_3$$

$$\rightarrow -1207 - ((-635) + (-394)) = \underline{\underline{178 \text{ kJmol}}}$$

Percentage error

$$\frac{\text{Standard Result} - \text{Our}}{\text{Standard Result}} \times 100 \quad \text{result}$$

$$\rightarrow \frac{178 - 126}{178} \times 100 = 29.21\% \\ = \underline{\underline{29\% \text{ to 2 sf}}}$$

Evaluation

Our experiment, when put into practice was not very reliable this was due to the factor of reactants surface area. When we used Calcium oxide chips the surface area was too small and by the time all the Calcium oxide had reacted the temperature had lowered not giving us an accurate maximum temperature. So we altered our experiment to use powdered calcium oxide so the reaction was quicker and more accurate.

Looking at my percentage error it is quite clear that the accuracy of my experiment was poor. This was due to many factors and sources of error: -

The major problem was that my experiment lost a lot of heat into the atmosphere, thermometer and into the glass beaker. This meant that our calculations would have been very inaccurate as they didn't account for the energy lost as heat and through other methods.

We could overcome this loss of energy by: using a polystyrene cup, this would insulate the liquid lowering the heat loss. Lowering the amount of air in the cup would lower the loss of heat into the atmosphere and as heat conventional current go upwards, a huge loss of heat would be through the top; so on adding a lid to the polystyrene cup, all this would lower our errors and raise our accuracy and reliability.

We cannot get any where near the accuracy of the Data book Enthalpy value because we could not simulate all the Standard conditions needed for our accurate results.

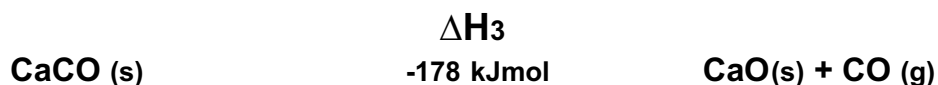
I also had problems with the equipment, the accuracy of my thermometer was low, as it only had graduations of 1°C. Where as a thermometer of graduations of 0.5 or 0.25°C would of produced a much more accurate reading for the temperature change. A minor error would be my glass wear error, because I used a measuring cylinder to measure my solutions, there would be a small percentage of liquid that would remain in the cylinder. I could lower this percentage by using a measuring cylinder with a higher grade quality of B or above would raise the accuracy and reliability of my results.

There was also a problem with the chemicals used, the calcium oxide, calcium carbonate and hydrochloric acid may have been contaminated with other chemicals. It might not all have reacted the way it should have. I could nullify this by using a brand new sample of chemicals and have no interactions with other experiments or uses. This would guarantee reliable results.

My method also had faults: The method did not specify a accurate measurement for Calcium carbonate and Calcium oxide, this would portray inaccurate and unreliable results, as an increase in the amount of reactants could increase the enthalpy value and could decrease the enthalpy value for a lower amount of reactants.

Conclusion

From my results I can tell that both reactions were exothermic, but that the reaction with calcium oxide was much more exothermic than the reaction with calcium carbonate. They were different because the bonds that were made in the calcium oxide reaction required less energy to be made than in the calcium carbonate reaction. My Hess' Law cycle can be labelled correctly:-



ΔH_1
- 1207 kJmol

ΔH_2
-1029 kJmol

CaCl (aq)

If I were to do the experiment again I would enforce my changes mentioned to achieve a closer result to the Standard enthalpy value. Also I would further investigate enthalpy changes of alcohol's.