Evaluation

I conducted an experiment to find out the enthalpy changes that took place for the reaction between copper sulphate and zinc. An enthalpy change of a reaction is the heat exchange with the surroundings at constant pressure.

Introduction

In this experiment in order to find out the enthalpy change I shall be reacting aqueous copper sulphate with zinc.

In this reaction Cu is being reduced and Zn is being oxidised. In this experiment the enthalpy change is exothermic.

Enthalpy changes can be both exothermic and endothermic. Exothermic reactions are easily recognised by a rise in temperature. Endothermic reactions require an energy input.

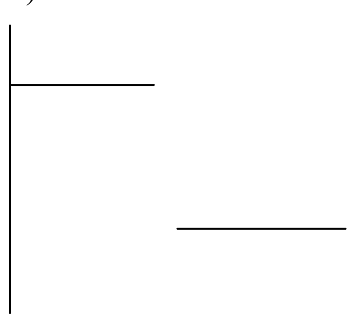
Enthalpy is the total energy content of the reacting materials it is given the symbol H. Enthalpy cannot be measured but it is possible to measure enthalpy change when energy is transferred to or from a reaction system. Enthalpy change is given the symbol ΔH where Δ is pronounced as change.

$\Delta H \quad = \quad H \text{ products} \quad - \quad H \text{ reactants.}$

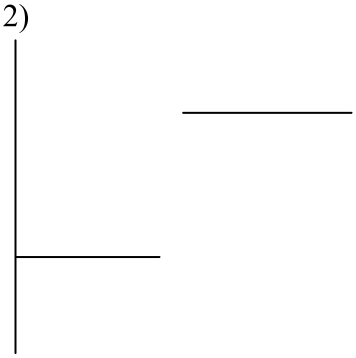
As ΔH is a measure of energy transferred to or from known amounts of reactants, the units are Kilojoules per mole

When we compare the enthalpy changes of various reactions we must use standard conditions such as known temperatures pressure amounts and concentration of reactants or products. This allows us to compare the standard enthalpy changes for reactions. A standard enthalpy for a reaction takes place under these standard conditions.
A pressure of 100 Kilopascals
A temperature of 298K
The reactants and products must be in the physical states (solid. liquid, or gas) that are normal for these conditions.

1)



2)



1) An exothermic enthalpy change is always given a negative value as the energy is lost from the system to the surroundings.
2) An endothermic enthalpy change is always given a positive value as the energy is gained by the system from the surroundings

Results

Time/ Min	0	0.5	1.0	1.5	2.0	2.5	3.0	3.5	4.0	4.5	5.0
Temperature ˚C	21	21	21	21	21	X	26	28	30	31	32

Time/Min	5.5	6.0	6.5	7.0	7.5	8.0	8.5	9.0	9.5	10.0	21.0
Temperature ˚C	32	33	34	34	34	35	35	35	36	36	30

Calculation

To calculate enthalpy change I would need the temperature change, the mass of copper(II)sulphate used and the specific heat capacity. For this experiment we must assume that the specific heat capacity of the solution is the same as water 4.18J
The temperature change is 16.5˚C, and the mass of copper(II) sulphate used is $50cm^3$, which is the same as 50g.

$$\Delta H = 4.18Jg \quad * \quad 50g \quad * \quad 16.5˚C$$
$$= 3448.5J$$

You then need to divide this by the no of moles of zinc used to get the overall enthalpy change.

$$\text{Moles} = \frac{\text{mass}}{\text{Ar}} = \frac{2.106}{65.4} = 0.032$$

$$\frac{3448.5}{0.032} = 107765$$

$$\frac{107765}{1000} = 107.77KJ$$

The overall enthalpy change given out by this experiment is 107.77KJ

Limitations and Improvements

There were two main limitations in this experiment, mainly the cup and the thermometer. The limitation of the cup was that it was not too well insulated so heat energy was lost more readily to the surroundings. Also there was no top so heat could have escaped through the top. The limitation with the thermometer was the range. By this I mean, when I took a temperature reading the closest I could get was 0.5, i.e. 21.5 31.5. If I had a better range on the thermometer I could get better, more accurate readings. Also when I took the temperature I could have touched the bottom and this may have affected the temperature.
Improvements I could have made in this experiment are I could have had a better insulated cup, this could have been done in two ways. I could have had two cups one inside the other this would have ensured that heat energy stayed in the cup longer, so my readings could be higher. The second is I could have had a copper cup. This would work as an insulator as it would absorb the heat so it would stay longer inside the solution. Also I could have had a top on my cup to prevent heat energy escaping from the top. Another improvement was too have a mechanical stirrer. The advantage of this is that I would then not have to use the thermometer as a stirrer. Also I could have had the thermometer in a fixed position so that my body heat would not interfere with the readings. Another improvement with the thermometer is that I could have had a larger range. If you look at my results you will se that the temperature varies between 20˚C and 40˚ C. If the thermometer was between these two temperatures and was more accurate then my results would also have been more accurate.
I could also use a data logger; this would take the temperature automatically and would also plot my points at regular intervals. This would help me as I would then not have to use the thermometer and I would not have made any errors. Also I could have had a colorimeter, which could have shown me when the reaction ended. This could have helped me because it would have been a point where I would have recognised when the temperature would start to fall. Another improvement is that I could have done the experiment again and gain another set of results to compare with my original results.

Errors and anomalous results.

The errors that occurred in this experiment are as follows. When I took a temperature reading my body heat may have interfered with the temperature. Also due to the lack of range on the thermometer I had to use my own guesswork as to what the temperature was when it was between two whole numbers. Also the experiment was not very efficient. By this I mean that a lot of energy was lost to the surrounding as there was a lack of insulation. If you look at my results you will see that there are not any anomalous results.