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ξ Class: 6S Class Number: 14 ξ
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Chemistry Full Report

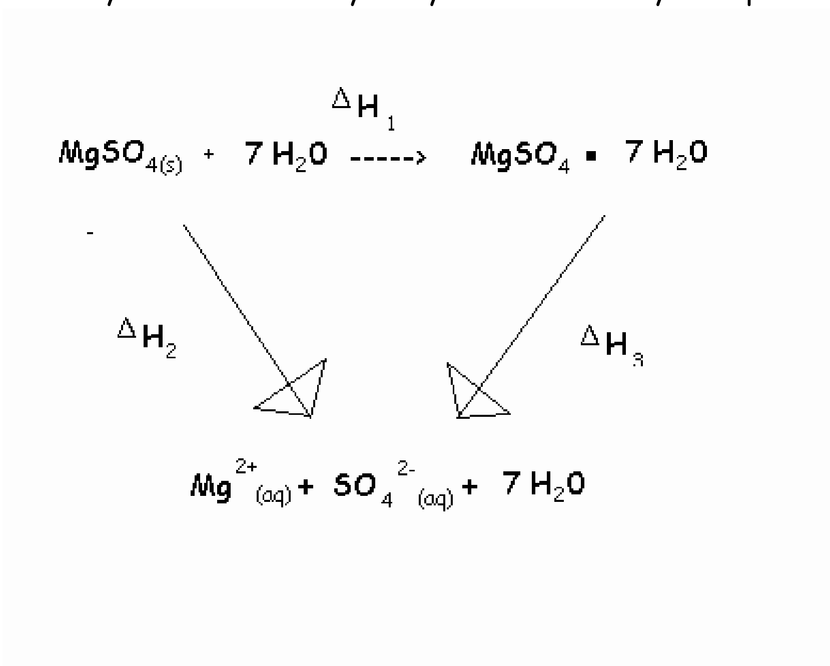
Experiment 5

Title: Application of Hess's Law to determine the enthalpy change of hydration of Magnesium sulphate (VI)

Aim: Using Hess's Law to find the enthalpy change of hydration of magnesium sulphate (VI)

Theory: Hess's Law can be defined as the heat given off or absorbed by a reaction is independent of the route taken.

In this experiment, the enthalpy change of hydration of Magnesium sulphate (VI) cannot be directly measured by calorimetry in the laboratory as hydration is a very slow process.



ΔH_1 = enthalpy change of hydration of $MgSO_{4(S)}$
ΔH_2 = molar enthalpy change of solution of hydrous $MgSO_{4(S)}$
ΔH_3 = molar enthalpy change of solution of $MgSO_4 \cdot 7H_2O_{(S)}$

According to Hess's law, the enthalpy of the overall reaction should be equal to the sum of the enthalpies of the two sub-reactions. Thus, this value can be calculated by applying Hess's Law

Procedure:

A. Determine the enthalpy change of solution of $MgSO_{4(S)}$

1. A balance was used to weigh the empty polystyrene foam cup
2. 50 cm^3 of deionized water was poured from the measuring cylinder to the polystyrene cup
3. The temperature of the water in the cup was measured using a thermometer
4. 0.025 mole of anhydrous magnesium sulphate (VI) was weighed accurately by the balance and was added into the foam cup
5. The solute was stirred to make sure all anhydrous magnesium sulphate (VI) were completely dissolved into the water as quickly as possible
6. The highest temperature of the solution was taken down
7. The molar enthalpy change of solution of $MgSO_{4(S)}$ was then calculated

B. Determine the enthalpy change of solution of $MgSO_4 \cdot 7H_2O_{(S)}$

1. 50 cm^3 of deionized water was poured from the measuring cylinder to the polystyrene cup again
2. Using a thermometer, the temperature of the water in the cup was measured
3. 0.025 mole of $MgSO_4 \cdot 7H_2O_{(S)}$ was weighed accurately by the balance instead of $MgSO_{4(S)}$and was added into the foam cup
4. The solute was stirred to make sure all anhydrous magnesium sulphate (VI) were completely dissolved into the water as quickly as possible
5. The highest temperature of the solution was taken down
6. The molar enthalpy change of solution of $MgSO_{4(S)}$ was then calculated

(Assume:
The specific heat capacity of the solution in the foam cup = $4.2kJkg^{-1}K^{-1}$
The specific heat capacity of the foam cup = $1.3kJkg^{-1}K^{-1}$)

Result:

Weight of the polystyrene cup = _________ g

	(A) Anhydrous	(B) Hydrated
Weight of magnesium sulphate added /g		
Initial temperature of water / C		
Highest temperature of the solution / C		
Lowest temperature of the solution / C		
Change in temperature / C		

Calculation:

1. The following assumptions were made:
   The specific heat capacity of the solution in the foam cup = $4.2 kJkg^{-1}K^{-1}$
   The specific heat capacity of the foam cup = $1.3 kJkg^{-1}K^{-1}$
   The thermal capacities of thermometer is negligible
   Density of the solution = $1 gcm^{-3}$
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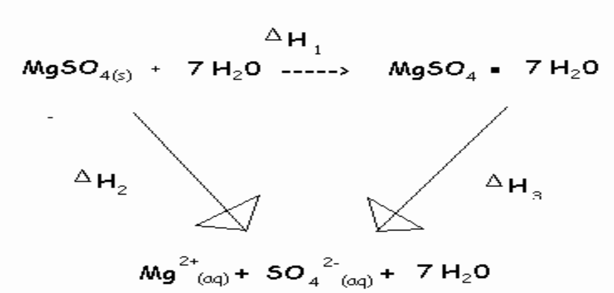
2. Calculate the molar enthalpy change of hydration of $MgSO_{4(S)}$
   ΔH for the reaction: $MgSO_{4(S)} + 7H_2O_{(S)} \rightarrow MgSO_4 \cdot 7H_2O_{(S)}$



Using Hess's Law ,

$$\Delta H1 = \Delta H2 - \Delta H3$$

$\Delta H2$:

$E = m\ c\ \Delta T$

$= [\ (\quad g\ )\ (\ 4.2 kJkg^{-1}K^{-1}\ ) + (\quad g\ )\ (\ 1.3 kJkg^{-1}K^{-1}\ )\ ]$

$[\ (\quad + 273\ ) - (\quad + 273\ )\ ]\ K$

$= \quad J$

Number of moles of $MgSO_4$ = 0.025 mol

$\Delta H2 = (\quad J\ ) / (\ 0.025\ mol\ )$

$= \quad kJ/mol$

molar enthalpy change of solution of hydrous $MgSO_{4(S)}$ = - kJ/mol

$\Delta H3$:

$E = m\ c\ \Delta T$

$= [\ (\quad g\ )\ (\ 4.2 kJkg^{-1}K^{-1}\ ) + (\quad g\ )\ (\ 1.3 kJkg^{-1}K^{-1}\ )\ ]$

$[\ (\quad + 273\ ) - (\quad + 273\ )\ ]\ K$

$= \quad J$

Number of moles of $MgSO_4$ = 0.025 mol

$\Delta H3 = (\quad J\ ) / (\ 0.025\ mol\ )$

$= \quad kJ/mol$

Molar enthalpy change of solution of $MgSO_4 \cdot 7H_2O_{(S)}$ = - kJ/mol

Using Hess's Law ,

$$\Delta H1 = \Delta H2 - \Delta H3$$

$= [\ (\qquad) - (\qquad)\ ]\ kJ/mol$

$= \qquad kJ/mol$

The true enthalpy change of the hydration of magnesium sulphate

= -104.0 kJ

Thus , the percentage error

= [(   ) – (   )] / (   )

=   %

Discussion:

3. Point out some factors leading to the difference between the experimental value and the true value

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4. Explain why the enthalpy change of the hydration of magnesium sulphate cannot be measured directly in the laboratory

The reason is that this reaction, hydration, is a very slow process which is

impossible to obtain directly in the laboratory

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5. Explain why it is not necessary to plot a temperature-time graph to determine the ΔT

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6. Draw an enthalpy level diagram for the reactions involved in the enthalpy cycle used