Chemistry Experiment Project

Class XII A

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Enthalpy of Formation of Magnesium Oxide

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1. Acknowledgements

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2. Introduction

The standard enthalpy of formation is an important quantity used by chemists, and also in the manufacturing industry. Often, it is very difficult to directly measure the enthalpy change in certain reactions. Hess's law is used in these cases to indirectly measure the enthalpy change of the required reaction by carrying out another series of reactions, and using the data gained by those reactions to compute the net enthalpy for the required reaction.

Magnesium Oxide is an important mineral that is used majorly for the manufacture of *Portland Cement*. Therefore, it is important to determine its properties, for it to be efficiently manufactured in industry conditions. Our objective in this experiment, is to measure one of the thermodynamic properties of MgO, its *standard enthalpy of formation*.

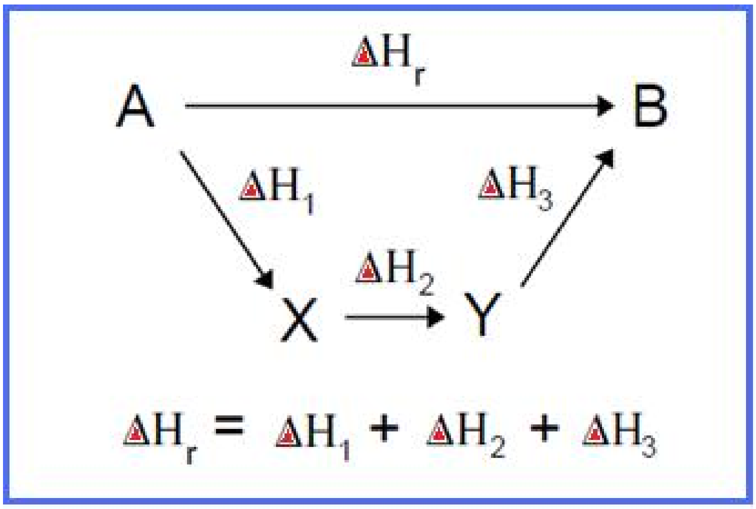
ΔH_r
A
B
ΔH_1
ΔH_3
ΔH_2
X
Y
$\Delta H_r = \Delta H_1 + \Delta H_2 + \Delta H_3$

Fig.1 Hess's Law

3. Experimental details

Enthalpy of formation of Magnesium Oxide

Aim

To determine the standard molar enthalpy of formation of Magnesium Oxide (MgO).

Materials Required

Calorimeter, thermometer, Bunsen Burner, Beaker, Tongs, Electric Balance.

Chemicals Required

Magnesium Ribbon, HCl (1.0 M), NaOH (1.0M)

Theory

The *standard enthalpy of formation* of a compound is the change of enthalpy during the formation of 1 mole of the substance from its constituent elements, with all substances in their standard states, and at a pressure of 1 bar.

The standard enthalpy of formation of Magnesium Oxide, would be the enthalpy of the reaction below:

$$\mathrm{Mg} + \frac{1}{2}\,\mathrm{O_2} \longrightarrow \mathrm{MgO} \tag{1}$$

Reaction (1) releases a lot of heat and light energy, and a direct method for measuring this energy would be extremely difficult in an experimental environment. Thus, we utilize Hess Law to indirectly measure the enthalpy of this reaction.

Hess' law states that the total enthalpy change during the complete course of a chemical reaction is the same whether the reaction is made in one step or in several steps.

Since it is easier and more accurate to measure the temperature change of a solution, we conduct the following reactions and measure the ΔT for each.

$$\mathrm{Mg\,(s) + 2\,H^+(aq) \longrightarrow Mg_2{}^+(aq) + H_2(g)} \tag{2}$$

$$\mathrm{MgO\,(s) + 2\,H^+(aq) \longrightarrow Mg_2{}^+(aq) + H_2O\,(l)} \tag{3}$$

$$\mathrm{\frac{1}{2}\,O_2(g) + H_2(g) \longrightarrow H_2O\,(l)} \tag{4}$$

We can clearly see that (2) - (3) + (4) = (1), thus the reactions can be used to determine the ΔH of formation of MgO. Note that reaction (4) is a standard reaction whose ΔH is known to be approximately 285.8 kJ/mol.

To measure ΔT for each reaction accurately, we conduct each reaction in a calorimeter, to prevent any loss of heat to the surroundings and thus minimize error. As reactions are carried out in a constant pressure environment, the heat released is equal to the enthalpy change in the reaction

In a solution of n moles, for a change in temp of ΔT, the heat released Q is given by:

$$Q = nC\Delta T \tag{5}$$

Where C is the molar heat capacity of the medium/body where change of temperature occurs.

In this experiment,we need to first determine the C of the calorimeter, as this value is unknown to us. For this, we conduct the following reaction separately in the calorimeter and measure the change in temperature ΔT_1:

$$\mathrm{nNaOH(1\,M) + nHCL(1\,M) \longrightarrow nNaCl + nH_2O} \tag{6}$$

The ΔH for this reaction is known to be -55.8 kJ/mol.
Also,

$$\text{Moles of water in the solution}(n') = \frac{2n}{18}$$

We also know that,

$$\text{Heat absorbed by calorimeter}(Q_{cal}) = \text{Heat released by reaction} - \text{Heat absorbed by Water}$$

$$\implies Q_{cal} = (-55800J \times n) - (\text{Moles of water in solution} \times 75.348J/molK \times \Delta T_1)$$

Thus, the heat capacity of the calorimeter is,

$$C_{cal} = \frac{Q_{cal}}{\Delta T_1} \tag{7}$$

After this value is known to use, we can use (5) to calculate the ΔH for reactions (1) to (4), and thus compute the enthalpy of formation of MgO.

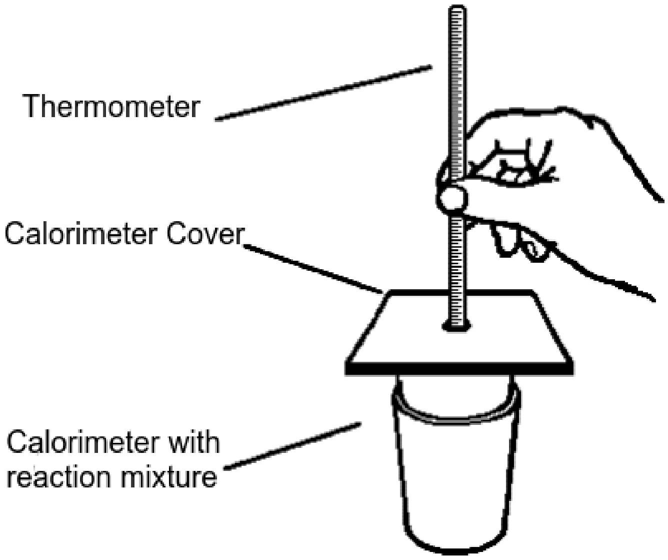
Thermometer
Calorimeter Cover
Calorimeter with reaction mixture

Fig.2 Experimental Setup for measuring temperature changes in reactions

Procedure

Determining Heat capacity of Calorimeter:

1. To determine heat capacity of calorimeter, the reaction between HCL and NaOH is carried out in aqueous solution in calorimeter.
2. Prepare standard solutions of 1M HCl by diluting 12M HCl and 1M NaOH by dissolving 40g NaOH pellets in 100 ml of water in standard flasks.
3. Take a calorimeter and place a thermometer in it. Note the initial reading of thermometer.
4. Add 15ml each of standard HCl and NaOH in the calorimeter.
5. Keep observing the thermometer until its temperature reaches a maximum value. Note the reading.
6. Calculate the value of heat capacity of the calorimeter.

Measuring enthalpy of formation of MgO:

1. Weigh 0.3g(0.025 mol) of Mg on an electronic balance.
2. Take the same calorimeter whose heat capacity was determined and add the 0.3g of Mg to it.
3. Take a thermometer and place it in the calorimeter. Note the initial reading of thermometer.
4. Add 25ml of standard HCl solution(1M) to the calorimeter. Note the maximum temperature of the thermometer.
5. Burn some Mg ribbons and take 0.5g(0.025 mol) of the MgO formed.
6. Wash the calorimeter and add the MgO to it.
7. Take a thermometer and place it in the calorimeter. Note the initial reading of thermometer.
8. Add 25ml of standard HCl solution(1M) to the calorimeter. Note the maximum temperature of the thermometer.
9. Use the obtained readings to calculate enthalpies of the above reactions per mole of HCl and hence determine the enthalpy of formation of MgO.

Observations

Temperature change for HCl + NaOH reaction (ΔT_1) = Final - Inital = $32.5 - 29 = 3.5°C$
Temperature change for Mg+HCl reaction (ΔT_2) = Final - Inital = $58 - 25 = 33°C$
Temperature change for MgO+HCl reaction (ΔT_3) = Final - Inital = $36 - 31 = 5°C$

Calculations

Calculating Heat Capacity of Calorimeter: By Equation (7) in Theory,

$$\begin{aligned} C_{cal} &= \frac{(55800J/molK * 0.015mol) - (\frac{30}{18}mol * 75.348J/molK * 3.5°C)}{3.5°C} \\ &= \frac{397.47}{3.5} \\ \implies C_{cal} &= 113.56J/K \end{aligned}$$

Since in the reactions of Mg with HCl and MgO with HCl, the solution had the same quantity of water (25 ml), thus we can include the heat absorbed by this water in our calorimeter's heat capacity, thus obtaining its *Effective Heat Capacity* (C_{eff}). Note that this is done purely for ease of calculation in subsequent steps.

$$\begin{aligned} C_{eff} &= (\frac{25}{18} \times 75.348) + 113.56 \\ &= 218.21J/K \end{aligned}$$

Thus, enthalpy change for Mg + HCl reaction (H_1) and for MgO + HCl (H_2) is given by:

$$\Delta H_1 = \frac{33 * 218.21}{0.025} = 288.03KJ/mol$$

$$\Delta H_2 = \frac{5 * 218.21}{0.025} = 43.642KJ/mol$$

Thus the enthalpy of formation of MgO within experimental errors is:

$$\begin{aligned} \Delta H_f &= 285.8 + 288.03 - 43.642 \\ &= 530.19kJ/mol \end{aligned}$$

Result

Known Standard Enthalpy of Formation of MgO is **601.8 kJ/mol**.
Calculated Standard Enthalpy of Formation of MgO is **530.2 kJ/mol**
Percentage Error = 12%

Precautions

1. Burn the magnesium ribbon carefully, holding the reactant with tongs away from face.

2. Handle Bunsen burner Carefully
3. Ensure to Cover Calorimeter to prevent heat loss
4. Avoid any parallax while taking temperature readings from the thermometer.
5. Wash all apparatus thouroughly before pouring in any reactant.

4. Conclusion

We conclude that Hess's law is a useful tool for indirectly measuring enthalpy changes for reactions, if the enthalpy changes are hard to directly measure. Another useful thing found from this experiment was that the reaction between NaOH and HCl can be used to easily determine the heat capacity of a calorimeter system.

5. Bibliography

Sources of Information:

- http://web.clark.edu/amixon/141_pm/f09formationmgo.pdf
- Chemistry Class XI Part 1 Book

We would also like to thank our chemistry teachers, Jayanti Ma'am and Lakshmi Ma'am for helping us with all the doubts we had during the course of this experiment.