

Hess's law lab essay sample

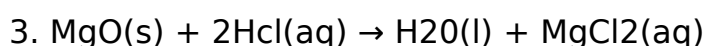
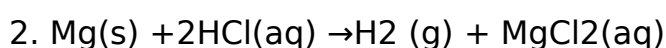


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Hess's law is a terminology in chemistry named after Germain Hess, a Swissborn Russian chemist and physician who published it in 1840. The law states that the total enthalpy change during the complete course of a reaction is the same whether the reaction is made in one step or in several steps. Enthalpy cannot be directly measured, but rather the change in enthalpy. Enthalpy is described as the change in kinetic and thermal energy. Hess's law states that the enthalpy change of a reaction will result in the same value regardless of what pathway is taken to achieve the products. In other words, only the start and end states matter to the reaction, not the individual steps between. This allows the change in enthalpy for a reaction to be calculated even when it cannot be measured directly. Chemical equations may be multiplied (or divided) by a whole number. When an equation is multiplied by a constant, its ΔH must be multiplied by the same number as well. If an equation is reversed, ΔH for the reaction must also be reversed. The addition of the equations can lead to the resulting net equation. If the net enthalpy change is negative, the reaction will be exothermic as it is releasing heat to the surrounding. If the net enthalpy change is positive, the reaction will be endothermic as it is absorbing heat from the surrounding.

Calorimetry is the scientific experimental measurement of the change of heat of an object or substance between its system and the surrounding. A calorimeter is a tool (usually a container) in which the heat exchange experiment is conducted, and is used to measure the quantity of heat transferred to or from the object. Heat exchange is the process that describes how the average molecular kinetic energy of a system is transferred to another system. Using proper terminology regarding heat

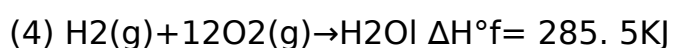
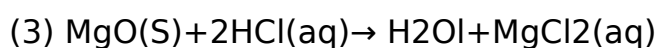
exchange from above, the equations can be deduced as $\Delta H = Q$ for endothermic reactions, and $\Delta H = +Q$ for exothermic reactions. Q in this case is the measure of the amount of heat released/absorbed, or rather the heat/energy gained by the surrounding and lost by the system, or vice versa. The equation to calculate the heat capacity is $Q = mc\Delta T$, where M is the mass, C is the specific heat capacity and Δt is the change in temperature. For the lab, four different reactions were listed. The net reaction, along with three other reactions was given: Net reaction: $\text{Mg(s)} + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{MgO(s)}$



4. $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2(\text{l})$ The enthalpy change for reaction two and three must be obtained experimentally, and the enthalpy change for reaction four was already given. To proceed, the first part of Hess's law must be applied to solve for the change in enthalpy of the net reaction using reaction two, three and four. Overall, the goal of this lab is to determine the enthalpy change in the net reaction by using the subsequent three reactions.

Purpose:

To determine the molar enthalpy change of the combustion of magnesium in the reaction $\text{Mg(s)} + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{MgO(s)}$ through experimentally obtaining the enthalpy change for reaction two and three listed below. Then, using the enthalpy change of reaction two, three and four and applying the first part of Hess's law to determine the enthalpy change of magnesium in the net equation. (2) $\text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{H}_2(\text{g}) + \text{MgCl}_2(\text{aq})$



Hypothesis:

The enthalpies of reaction two and three can be found experimentally by measuring the change in temperature of the reaction and using the equation: $mc\Delta T$. The mass can be determined by using the volume of HCl used and since hydrochloric acid is almost the same density as water which is 1 g/cm^3 or 1 g/mL . Since 1 mol/L hydrochloric acid is mostly water, the heat capacity used would be the same as water's which is $4.19 \text{ J/g}^\circ\text{C}$. The enthalpy of reaction two needs to be multiplied by 1 because the equation needs to be flipped because Magnesium oxide needs to be on the product side of the equation. The enthalpy change for reaction two and four will remain as they are once they have been calculated, as it does not need to be flipped or multiplied/divided. Finally when all of the enthalpies of the reaction have been arithmetically equated, it will equal to the molar enthalpy of combustion for magnesium. It should be a negative number because combustion is an exothermic reaction that releases heat to the surroundings. Procedure:

- 1) The mass of a magnesium strip was measured and recorded at 0.36 g using the scale.
- 2) Hydrochloric acid was acquired and poured into the graduated cylinder. The mass of $\text{HCl}(aq)$ was measured and recorded at 89.90 mL using the scale.
- 3) A closed calorimeter was created using a Styrofoam cup and cardboard cup with a hole, stacking the cardboard cup on top of the Styrofoam cup.
- 4) Hydrochloric acid was poured into the calorimeter, a thermometer was inserted through the hole and the initial temperature of

the HCl(aq) was measured at 24°C. 5) The magnesium strip was inserted into the Hydrochloric acid filled calorimeter for approximately five minutes. 6) After reaction completed, the final temperature of the Hydrochloric acid was recorded with the thermometer at 41. 2°C. 7) The mass of magnesium oxide was acquired. The mass was measured and recorded at 0. 77g using the scale. 8) Hydrochloric acid was acquired and poured into the graduated cylinder. The mass of HCl(aq) was measured and recorded at 90. 06mL using the scale. 9) The closed calorimeter was recreated by stacking a Styrofoam cup and cardboard cup with a hole. 10) Hydrochloric acid was poured into the calorimeter, a thermometer was inserted through the hole and the initial temperature of the HCl(aq) was measured at 24°C. 11) The magnesium oxide was inserted into the Hydrochloric acid filled calorimeter for approximately five minutes. 12) After reaction completed, the final temperature of the Hydrochloric acid was recorded at 29°C. 13) The hydrochloric acid was disposed of safely, and the work station was cleaned.

Observations:

Solute used| MgO(s)| Mg(s)|

Mass| 0. 77 g| 0. 36 g|

Volume of HCl(aq)| 90. 06 mL| 89. 9 mL|

Concentration of HCl(aq)| 1 mol/L| 1 mol/L|

Initial temperature| 24. 0°C| 24. 0°C|

Final temperature| 29. 0°C| 41. 2°C|

Quantitative observation chart for reaction two and three:

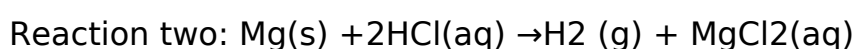
Qualitative observation chart for reaction two and three:

Solute/Solvent| MgO (s)| Mg(s)| HCl(aq)|

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Before the reaction:| White - coloured, solid at room temperature.| A thin piece of metal, greyish colour. Solid at room temperature.| A clear, colourless solution that is in water. | After the reaction:| None remaining; all dissolved in the Hydrochloric acid. Therefore it is the limiting reactant.| None remaining; all dissolved in the Hydrochloric acid. Therefore it is the limiting reactant.| Remained the same.| Potential Energy Diagram:

Analysis:



Limiting reactant:



$0.015\text{mol} | 0.015(2) = 0.03\text{mol}$

$2.46(1/2) = 1.23\text{mol} | 2.46\text{mol}$

Therefore Mg is the limiting reactant as HCl is in excess.

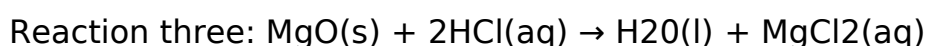
$q_{\text{rxn}} = q_{\text{water}} = n_{\text{Mg}} \cdot m_{\text{M}} = MC\Delta T = 0.46\text{g} \cdot 24.31\text{g/mol} = (89.9\text{g})(4.19\text{J/g}^\circ\text{C})$

$(41.0^\circ\text{C} - 24.0^\circ\text{C}) = 0.015\text{mols} = 6.42\text{KJ}$

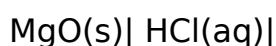
$\Delta H^\circ = \Delta H_n$

$= 6.42\text{KJ} / 0.015\text{mols}$

$= 426.70\text{KJ/mol}$



Limiting reactant:



$0.019\text{mol} | 0.019(2) = 0.038\text{mol}$

$2.47(1/2) = 1.24\text{mol} | 2.47\text{mol}$

Therefore MgO is the limiting reactant as HCl is in excess.

$$q_{\text{rxn}} = q_{\text{water}} = n_{\text{MgO}} \cdot M = MC\Delta T = 0.019 \text{ mol} \cdot 24.31 + 16 = (90.1) (4.19 \text{ J/g}^\circ\text{C}) \\ (29.0^\circ\text{C} - 24.0^\circ\text{C}) = 1.90 \text{ kJ}$$

$$\Delta H^\circ = \Delta H_n$$

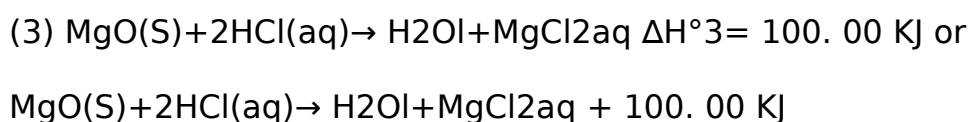
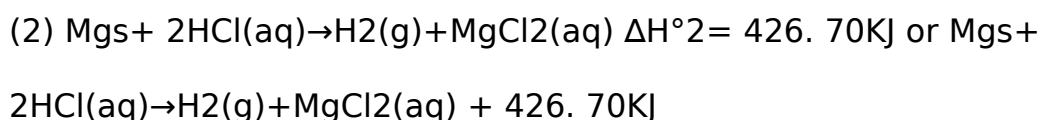
$$= 1.90 \text{ kJ} / 0.019 \text{ mol}$$

$$= 100.00 \text{ kJ/mol}$$

Assumptions:

The first major assumption was that no heat was lost to the surroundings other than to the water. It was assumed that the system was a completely isolated system that does not release or absorb heat or matter from the surrounding. This is not necessarily true as the calorimeter was not very insulated so heat could escape and there was a hole in the cup's lid that allowed the mass to be released, so it was in reality an open system. The second major assumption was to use the heat capacity of water to do the equation: $mc\Delta T$ when in reality the specific heat capacity was in regards to the hydrochloric acid, therefore it is not exactly $4.19 \text{ J/g}^\circ\text{C}$. Though the solution was mostly water, heat capacity would be different than that of water's because there is hydrochloric acid in the water that are going through the temperature change and absorbing energy along with the water.

Thermochemical equation for reaction two and three:



Molar enthalpy of combustion for magnesium:

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The final net equation will be: $\text{Mg(s)} + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{MgO(s)}$

$\Delta H^\circ_{\text{comb}}(\text{Mg}) = \Delta H^\circ_2 + \Delta H^\circ_3 + \Delta H^\circ_4$ $\Delta H^\circ_3 \times 1$ because equation is flipped

$= 426.70 \text{ kJ/mol} + (1)100 \text{ kJ/mol} + (285.5 \text{ kJ/mol})$

$= 612.50 \text{ kJ/mol}$

Accepted value:

Accepted value $\Delta H^\circ_{\text{comb}}(\text{Mg}) = 601.83 \text{ kJ/mol}$

Percent error = $\frac{\text{experimental value} - \text{accepted value}}{\text{accepted value}} \times 100\%$

$= \frac{(612.50 \text{ kJ/mol}) - (601.83 \text{ kJ/mol})}{(601.83 \text{ kJ/mol})} \times 100\%$

$= 1.77\%$

Sources of error:

The first major source of error was the loss of heat and energy to the surroundings that is not accounted for. The calorimeter used was not an isolated system but rather an open system, heat could have escaped through or even absorbed by the Styrofoam cup it. Heat and matter can be released and/or absorbed by the system as there was a hole in the top of the cup that could not be completely insulated. This deflates the actual results and $\Delta H^\circ_{\text{rxn}}$ should have been greater, according to the equation $\Delta H^\circ_{\text{rxn}} = q_{\text{surroundings}}$. Theoretically, the heat gained should not solely account for the water, as the heat gained between the air etc. must also be taken into consideration. Therefore, ΔH should be greater. To improve this, the size of the hole on the cup can be made accordingly to the thermometer's width, which the purpose is to let less matter and heat escape. Also, using more advanced calorimeters (As opposed to plastic cups) that unfortunately the school's science department cannot afford could also improve the accuracy

and efficiency of the result. An example is a bomb calorimeter, which creates an approximate isolated system so there would be significantly less heat escaping to the surroundings that are not accounted for.

The other major source of error comes from the problem when transferring the reactant. Firstly, when transferring hydrochloric acid from the graduated cylinder to the Styrofoam cup, not all of it was transported as some stuck to the cylinder, this was unpreventable as it is physically impossible to reach into the graduated cylinder and pull out the droplets of the hydrochloric acid. This inflated the ΔH as the equation used is $\Delta H = MC\Delta T$, both the heat change and enthalpy are correlated, as the mass increases so will ΔH . Therefore, the volume of the hydrochloric acid used was actually less than the volume recorded. A solution is using a calorimeter that has marking for volume already on it so hydrochloric acid can be poured directly into it, therefore there would be an accurate reading of the HCl(aq) solution used in regards to its mass, as opposed to losing a small portion of its volume when transferring the reactant.

Lastly, there was a problem with the magnesium strip as it quickly reacts with the oxygen in the air to form MgO(s) during the timeframe when the magnesium was obtained and when it was used in the reaction. This would deflate the molar enthalpy because according to the equation $n\text{Mg} = m\text{M}$, as the mass of the magnesium will react with the oxygen in the air and the mass will increase which will consequently also increase the number of moles, therefore when looking at the equation $\Delta H^\circ = \Delta H_n$, the higher the number of moles, the lower the molar enthalpy will be. Therefore, it's essentially recorded as more moles than what it actually is because some of

it reacted to form magnesium oxide, the molar enthalpy of the recorded mass would be higher than in reality. This error can be avoided by having the scale, magnesium and calorimeter next to each other. The Mg(S) can be quickly weighed after cleaned and dropped immediately into the calorimeter instead of traveling from the front of class to the back, and therefore avoid the oxidation of the magnesium. t in. Conclusion:

In conclusion this lab was successful in using Hess' law to determine the molar enthalpy of combustion for magnesium through experiments. The hypothesis was correct as the enthalpy was in fact negative. The experimental value is = 612. 50KJ/mol and the accepted value is 601. 83KJ/mol which only equates to a percentage error of 1. 77%. The results were fairly accurate but can be improved by avoiding the sources of error listed above.

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