

The following lab based experiment will serve to test your knowledge of thermochemistry ...

can you stand the heat? cue the music ...

Part A:

A constant pressure calorimeter was assembled. It will be used in parts B & C. Before parts B & C can be carried out, we must determine the heat capacity of the calorimeter. Cold water is placed in a calorimeter and warm water will be added. The resultant equilibrium temperature will be measured and the C_{cal} can be calculated.

Cold water

30.00 grams

20.00°C (initial temp)

Warm water

26.52 grams

47.00°C (initial temp)

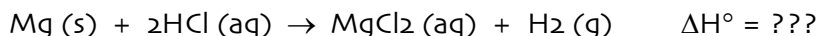
Equilibrium data

31.00°C (final temp of the cal. & contents)

1) What is the heat capacity (C_{cal}) of the calorimeter?

Part B:

The following reaction will be carried out in the calorimeter prepared in Part A.



100.00 mL of 1.000M HCl were added to the calorimeter. After resting a while the temperature of the calorimeter and its contents were measured to be 20.00°C.

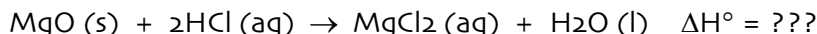
To the calorimeter 0.03647 grams of magnesium metal are added. The reagents are swirled until reaching a maximum temperature of 21.61°C. Assume that the density of the resulting solution (which is mostly HCl) is 1.020 g/mL and that its specific heat is 3.91 J/gK.

2) Was there a limiting reagent? If so, what reactant was depleted first?

3) What is the ΔH° for this reaction?

Part C:

The following reaction will be carried out in the calorimeter prepared in Part A.



100.00 mL of 1.000M HCl were added to the calorimeter. After resting a while the temperature of the calorimeter and its contents were measured to be 20.00°C.

To the calorimeter 1.209 grams of magnesium oxide powder are added. The reagents are swirled until reaching a maximum temperature of 30.42°C. Assume that the density of the resulting solution (which is mostly HCl) is 1.020 g/mL and that its specific heat is 3.91 J/gK.

4) Was there a limiting reagent? If so, what reactant was depleted first?

5) What is the ΔH° for this reaction?

Part D:

Take the experimentally derived results from Parts B & C and apply Hess' Law to determine the ΔH_f° for MgO. Remember the demo for this rxn? It was quite energetic and I mentioned that you should not stare directly at the flame. Now you know why we cannot experimentally determine this value using a constant pressure calorimeter. Have fun!

6) What is the ΔH_f° for MgO?

7) How could you find accepted values for this experiment?

Part E:

Use the summation law ($\Delta H_{\text{rxn}} = \sum n\Delta H_f^\circ \text{ products} - \sum m\Delta H_f^\circ \text{ reactants}$) and Hess' Law (if needed) to find the ΔH_f° for MgCl_2 and the ΔH_f° Mg^{2+} .

8) What is ΔH_f° for MgCl_2 ?

9) What is ΔH_f° for Mg^{2+} ?

ANSWERS:

PART A:

$$\Delta T \text{ hot water} = 31.00^\circ\text{C} - 47.00^\circ\text{C} = -16.00^\circ\text{C} \quad \Delta T \text{ cold water} = 31.00^\circ\text{C} - 21.00^\circ\text{C} = 11.00^\circ\text{C}$$

$$26.52\text{g} (4.184\text{J/gK}) -16.00^\circ\text{C} = [(30.00\text{g} \times 4.184\text{J/gK} \times 11.00^\circ\text{C}) + (C_{\text{cal}} \times 11.00^\circ\text{C})]$$

$$-1775\text{J} = -(1381\text{J} + C_{\text{cal}} \times 11.00^\circ\text{C}) \quad -1775\text{J} = -1381\text{J} - C_{\text{cal}} \times 11.00^\circ\text{C}$$

$$-394\text{J} = -C_{\text{cal}} \times 11.00^\circ\text{C} \quad 1) -394\text{J}/-11.00^\circ\text{C} = 35.8\text{J}/^\circ\text{C}$$

PART B:

$$\Delta T = 21.61^\circ\text{C} - 20.00^\circ\text{C} = 1.61^\circ\text{C}$$

$$\Delta H_{\text{rxn}} = -[(100.00\text{g} \times 1.020\text{g/ml} \times 3.91\text{J/g}^\circ\text{C} \times 1.61^\circ\text{C}) + (35.8\text{J}/^\circ\text{C} \times 1.61^\circ\text{C})]$$

$$\Delta H_{\text{rxn}} = -(642\text{J} + 57.6\text{J}) \quad \Delta H_{\text{rxn}} = -700\text{J}$$

2) The limit is Mg!

$$3) \Delta H^\circ = \text{kJ/mol (need moles of the L.R.)} \quad 0.03647\text{g Mg} \times 1\text{ mol Mg}/24.31\text{ g} = 0.001500\text{ mol Mg}$$

$$\text{So, } -700\text{J}/0.001500\text{ mol} \times 1\text{kJ}/100\text{J} = -467\text{ kJ/mol Mg}$$

PART C:

$$\Delta T = 30.42^\circ\text{C} - 20.00^\circ\text{C} = 10.42^\circ\text{C}$$

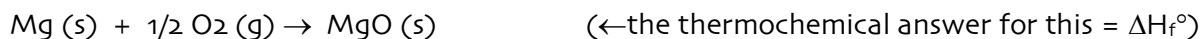
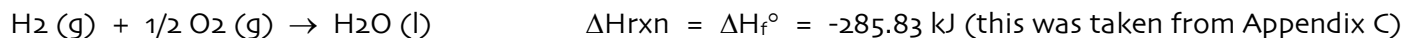
$$\Delta H_{\text{rxn}} = -[(100.00\text{g} \times 1.020\text{g/ml} \times 3.91\text{J/g}^\circ\text{C} \times 10.42^\circ\text{C}) + (35.8\text{J}/^\circ\text{C} \times 10.42^\circ\text{C})]$$

$$\Delta H_{\text{rxn}} = -(4160\text{J} + 373\text{J}) \quad \Delta H_{\text{rxn}} = -4530\text{J}$$

4) The limit is MgO!

$$5) 1.209\text{ g MgO} \times 1\text{ mol MgO}/40.31\text{g MgO} = 0.03000\text{ mol MgO}$$

$$-4530\text{J}/0.03000\text{ mol MgO} \times 1\text{ kJ}/1000\text{J} = -151\text{ kJ/mol MgO}$$

PART D:

$$6) \begin{array}{r} +151\text{ kJ} \\ -467\text{ kJ} \\ \hline -285.83\text{ kJ} \\ \hline -602\text{ kJ} \end{array} = \Delta H_f^\circ \text{ (for MgO formation)}$$

7) Use appendix C and the summation law (Σ). If ΔH_f° values exist for the reagents and products, any thermochemical equation can be directly calculated.

PART E:



$$\Delta H_{\text{rxn}} = \sum n\Delta H_f^\circ \text{ products} - \sum m\Delta H_f^\circ \text{ reactants} \quad -467 \text{ kJ} = [\Delta H_f^\circ \text{MgCl}_2 \text{ (aq)}] - [2(\Delta H_f^\circ \text{HCl (aq)})]$$

$$-467 \text{ kJ} = \Delta H_f^\circ \text{MgCl}_2 \text{ (aq)} - [2(-167.2 \text{ kJ})] \quad = \quad -467 \text{ kJ} = \Delta H_f^\circ \text{MgCl}_2 \text{ (aq)} + 334.2 \text{ kJ}$$

$$8) \quad -801. \text{ kJ} = \Delta H_f^\circ \text{MgCl}_2 \text{ (aq)}$$

9) MgCl (aq) is really just $\text{Mg}^{2+} \text{ (aq)}$ and $2\text{Cl}^- \text{ (aq)}$ because MgCl_2 dissociates 100% in water.

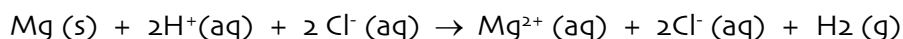
From Appendix C:

$$\Delta H_f^\circ \text{HCl (aq)} = -167.2 \text{ kJ}$$

Look at $\Delta H_f^\circ \text{H}^+ \text{ (aq)}$. It is 0.00 kJ! Therefore, $\Delta H_f^\circ \text{HCl (aq)} = \Delta H_f^\circ \text{Cl}^-$

This is because HCl (aq) is really $\text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$

So, $\text{Mg (s)} + 2\text{HCl (aq)} \rightarrow \text{MgCl}_2 \text{ (aq)} + \text{H}_2 \text{ (g)} \quad \Delta H_{\text{rxn}} = -467 \text{ kJ}$ or you could see it as



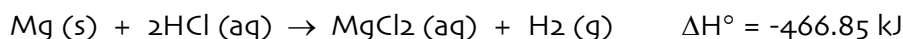
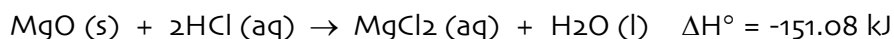
. . . a net ionic equation . . . $\text{Mg (s)} + 2\text{H}^+ \text{ (aq)} \rightarrow \text{Mg}^{2+} \text{ (aq)} + \text{H}_2 \text{ (g)}$

$$\Delta H_{\text{rxn}} = \sum n\Delta H_f^\circ \text{ products} - \sum m\Delta H_f^\circ \text{ reactants}$$

$$-467. \text{ kJ} = [(\Delta H_f^\circ \text{Mg}^{2+} \text{ (aq)})] \quad \text{all other reagents are elemental or } \text{H}^+ \text{ (which is also zero)!}$$

$$\Delta H_f^\circ \text{Mg}^{2+} \text{ (aq)} = -467. \text{ kJ}$$

ACCEPTED VALUES (FROM APPENDIX C AND DERIVED)



$$\Delta H_f^\circ \text{MgO (s)} = -601.6 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{H}_2\text{O (l)} = -285.83 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{HCl (aq)} = -167.2 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{H}^+ \text{ (aq)} = 0 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{MgCl}_2 \text{ (aq)} = -801.25 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{Mg}^{2+} \text{ (aq)} = -466.85 \text{ kJ/mol}$$