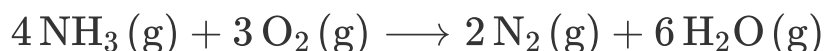


HL Chemistry - Unit 5 Test: Energetics (2020)

1. The enthalpy of reaction for the following reaction can be determined theoretically or experimentally.



a. An equation could be used to determine the enthalpy change of this reaction using bond enthalpies. Without using any numerical values, show how the enthalpy change would be calculated by this method

The enthalpy value can be calculated by subtracting the sum of the bond dissociation energy of the reactants by the sum of the bond dissociation energy of the products.

$$\begin{aligned}\Delta H &= \sum \text{BDE} (\text{reactants}) - \sum \text{BDE} (\text{products}) \\ &= [4 \times 3 \times (\text{N-H}) + 3 (\text{O=O})] - [2 (\text{N}\equiv\text{N}) + 6 \times 2 \times (\text{H-O})]\end{aligned}$$

Define average bond enthalpy.

Average bond enthalpy is the average energy required to break/form a bond between two elements in gaseous state. The value is not accurate as bond enthalpy changes in different molecular geometries.

2. The enthalpy of this reaction can also be determined experimentally in a school laboratory. Stoichiometrically matched quantities of ammonia and oxygen were weighed and placed in an insulated bomb calorimeter containing 80.0 mL \pm 0.5 water as illustrated below.

$$M_{\text{NH}_3}(\text{g}) = 0.24 \pm 0.01$$

$$M_{\text{O}_2}(\text{g}) = 0.34 \pm 0.01$$

$$T_{\text{initial}}(^{\circ}\text{C}) = 19.05 \pm 0.02$$

$$T_{\text{final}}(^{\circ}\text{C}) = 8.74 \pm 0.02$$

a. Using the data available, calculate the change in temperature with the correct level of precision.

$$\Delta T = T_{\text{final}} - T_{\text{initial}}$$

$$= 8.74 \pm 0.02 - 19.05 \pm 0.02$$

$$= -10.31 \pm 0.04^{\circ}\text{C}$$

b. Use the data to calculate the heat absorbed or released (in kJ) when the reaction occurred.

$$q = mc\Delta T$$

$$= 80.0 \text{ mL} \times 1 \text{ g/mL} \times 4.18 \text{ J/g}^{\circ}\text{C} \times (-10.31^{\circ}\text{C})$$

$$= -1723.8 \text{ J}$$

$$= -1.72 \text{ kJ}$$

$$\Delta H = -q$$

$$= 1.72 \text{ kJ}$$

Therefore the reaction absorbs 1.72 kJ of heat.

c. Calculate the enthalpy change per mole of $\text{NH}_3(\text{g})$ reacted.

$$\begin{aligned}
 M_{\text{NH}_3} &= M_{\text{N}} + 3 \times M_{\text{H}} \\
 &= 14.01 + 3 \times 1.01 \\
 &= 17.04 \text{ g/mol}
 \end{aligned}$$

$$0.24 \text{ g} / 17.04 \text{ g/mol} = 0.014 \text{ mol}$$

$$0.82 \text{ kJ} / 0.014 \text{ mol} = 58.22 \text{ kJ/mol}$$

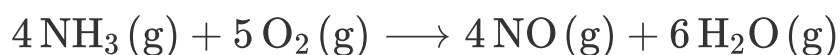
d. The final temperature reading in the data above was actually obtained by collecting data over a 10 minute period. Sketch a graph showing the data collected and best method for determining the final temperature to be used in a calorimetry experiment.

The temperature drop rate can be measured and the data can be extrapolated to determine what the final temperature would have been if there was no heat loss to the environment. This can be determined using a graphing software or simply drawing a line through the temperature drop.

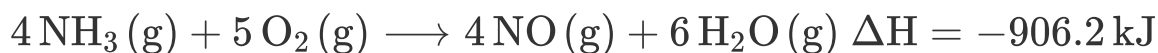
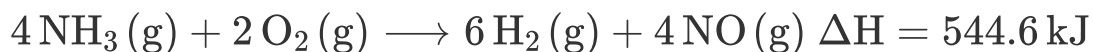
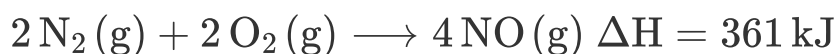
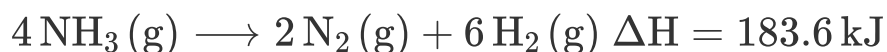
e. Describe one potential source of systematic error for this experiment that is NOT corrected for in the graph above.

One source of potential error that is not corrected by the temperature extrapolation would be that the specific heat capacity of the bomb calorimeter has not been considered and this will result in a lower final temperature of the solution.

3. Using the intermediate reactions provided, determine the molar enthalpy for the reaction below

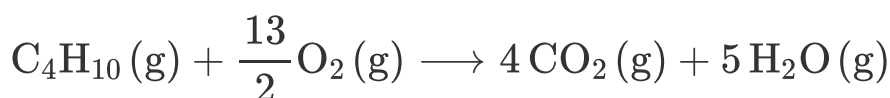


This problem can be solved by implementing Hess' Law:



Therefore the enthalpy of the reaction is $\Delta H = -906.2 \text{ kJ}$

4. Determine the standard molar enthalpy for complete combustion for butane (C₄H₁₀) using the values in table 12.



$$\Delta H_{\text{rxn}} = \sum \text{BDE}(\text{reactants}) - \sum \text{BDE}(\text{products})$$

$$\sum \text{BDE}(\text{reactants}) = 3 \times (\text{C}-\text{C}) + 10 \times (\text{C}-\text{H}) + \frac{13}{2} \times (\text{O}=\text{O})$$

$$= 3 \times 346 \text{ kJ} + 10 \times 414 \text{ kJ} + \frac{13}{2} \times 498 \text{ kJ}$$

$$= 8415 \text{ kJ}$$

$$\sum \text{BDE}(\text{products}) = 4 \times 2 \times (\text{C}=\text{O}) + 5 \times 2 \times (\text{H}-\text{O})$$

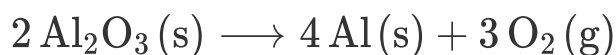
$$= 8 \times 804 \text{ kJ} + 10 \times 463 \text{ kJ}$$

$$= 11062 \text{ kJ}$$

$$\Delta H = 8415 \text{ kJ} - 11062 \text{ kJ}$$

$$= -2647 \text{ kJ}$$

5. When aluminum ore is smelted, aluminum is separated according to the equation:



Use the data below to answer the questions that follow.

$$S_{2\text{Al}_2\text{O}_3}^{\circ} = 50.9 \text{ J/K} \times \text{mol}$$

$$S_{\text{O}_2}^{\circ} = 205.0 \text{ J/K} \times \text{mol}$$

$$S_{\text{Al}}^{\circ} = 28.3 \text{ J/K} \times \text{mol}$$

a. Calculate the standard entropy change for the reaction.

$$\begin{aligned}\Delta S &= \sum S_{\text{product}} - \sum S_{\text{reactant}} \\ &= (4 \times 28.3 \text{ J/K} \times \text{mol} + 3 \times 205.0 \text{ J/K} \times \text{mol}) - 50.9 \text{ J/K} \times \text{mol} \\ &= 677.3 \text{ J/K} \times \text{mol}\end{aligned}$$

b. If ΔH° for the reaction is -138 kJ Calculate a value for ΔG° for the reaction and explain what can be inferred from your answer.

$$\begin{aligned}\Delta G &= \Delta H - T\Delta S \\ &= -138 \text{ kJ} - 298 \text{ K} \times 677.3 \text{ J/K} \times \text{mol} \\ &= -340 \text{ kJ}\end{aligned}$$

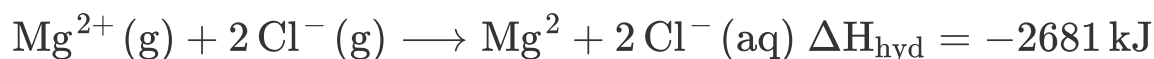
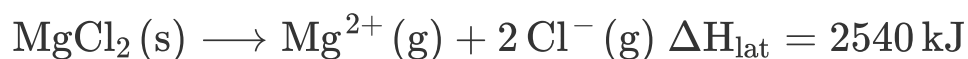
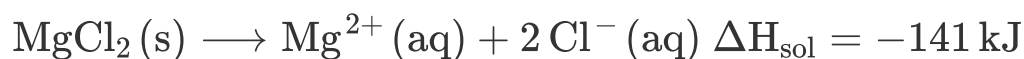
Because the ΔH value of the reaction is negative and the ΔS value is positive, therefore the ΔG value will always be negative and the reaction will be spontaneous at all temperatures.

c. State what change in conditions would be required to produce the opposite interpretation for your answer in (b).

Because both the reaction is always spontaneous, no change to the temperature can yield a change to the spontaneity of the reaction.

6. Draw a fully labeled enthalpy (energy) cycle for the dissolving of magnesium chloride, including values where appropriate.

a. Draw a fully labeled enthalpy (energy) cycle for the dissolving of magnesium chloride



b. Use your enthalpy cycle in part (a) to determine the enthalpy of solution of magnesium chloride.

$$\begin{aligned}\Delta H_{\text{sol}} &= \Delta H_{\text{lat}} + \Delta H_{\text{hyd}} \\ &= 2540 \text{ kJ} + (-2681 \text{ kJ}) \\ &= -141 \text{ kJ}\end{aligned}$$

c. One of the values used to calculate enthalpy of solution is lattice enthalpy. Explain why magnesium sulfide has a higher lattice enthalpy than magnesium chloride.

Magnesium Sulfide has a higher lattice enthalpy than Magnesium Chloride because MgS forms $2+$ and $2-$ ions whereas MgCl_2 forms $2+$ and $1-$ ions. As lattice enthalpy is proportional to $\frac{\text{charge}}{\text{radius}}$, an increasing in charge will result in an increase in ΔH_{lat} .