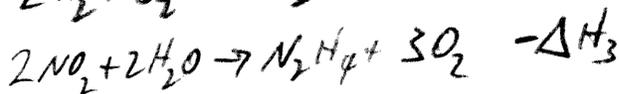
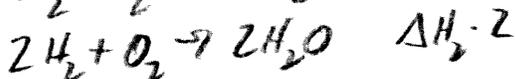
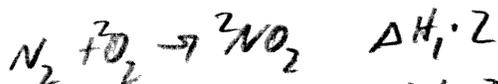
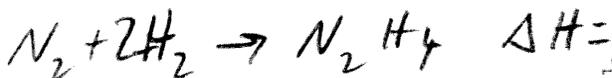
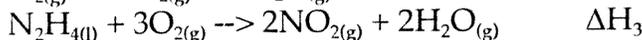
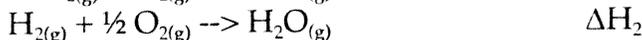
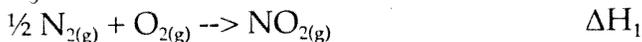


More Ch.6 AP Questions

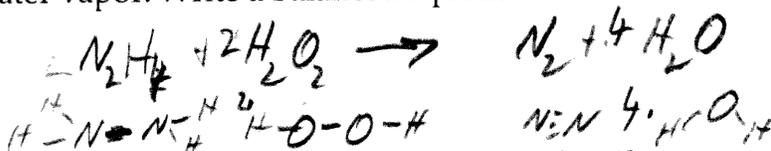
1. Liquid hydrazine, N_2H_4 , is sometimes used as a rocket propellant.

a. Write an equation for the formation of hydrazine from its elements and use the combustion equations below to derive an equation in which ΔH_f° for liquid hydrazine is expressed in terms of ΔH_1 , ΔH_2 and ΔH_3 .



$$\Delta H_f^\circ = (2 \cdot \Delta H_1) + (2 \cdot \Delta H_2) - \Delta H_3$$

b. In a rocket, liquid hydrazine is reacted with liquid hydrogen peroxide to produce nitrogen and water vapor. Write a balanced equation for this reaction.



c. Calculate ΔH°_{rxn} for the reaction represented in 2b.

	ΔH_f° kJ/mol
$N_2H_{4(l)}$	50.6
$H_2O_{2(l)}$	-187.8
$H_2O_{(g)}$	-285.8

$$\Delta H = [\Delta H_{N_2} + 4\Delta H_{H_2O}] - [\Delta H_{N_2H_4} + 2\Delta H_{H_2O_2}]$$

$$\Delta H = [0 + 4 \cdot (-285.8)] - [50.6 + 2 \cdot (-187.8)]$$

$$\Delta H = -1,143.2 - (-325) = -818.2 \frac{kJ}{mol}$$

d. Calculate ΔH°_{rxn} for the reaction in 2b using bond dissociation energies.

Bond Dissociation Energy kJ/mol	
N-N 167	O-O 142
N=N 418	O=O 494
N≡N 942	O-H 459
N-H 386	

$$\Delta H = [4 \cdot N-H + N-N + 4 \cdot H-O + 2O-O] - [N \equiv N + 4 \cdot O-H]$$

$$\Delta H = [4 \cdot 386 + 167 + 2 \cdot 142] - [942 + 4 \cdot 459]$$

$$1,995 - 2,778$$

$$\Delta H = -783 \frac{kJ}{mol}$$

e. Which value of $\Delta H^\circ_{\text{rxn}}$ (that calculated in part c or part d) is likely to be more accurate? Justify your answer.

ΔH_f values are probably more accurate they take into account the influence of the whole molecule

f. Calculate the maximum temperature of the combustion gases if all the energy generated in the reaction goes into raising the temperature of those gases. [The heat capacities of $\text{N}_2(\text{g})$ and $\text{H}_2\text{O}(\text{g})$ are $29.1 \text{ J}/(\text{mol}\cdot^\circ\text{C})$ and $33.6 \text{ J}/(\text{mol}\cdot^\circ\text{C})$, respectively.]

$$q = -818.2 \frac{\text{kJ}}{\text{mol}} \quad \text{assume } 1 \text{ mol } \text{N}_2\text{H}_4 \rightarrow 1 \text{ mol } \text{N}_2 \\ 4 \text{ mol } \text{H}_2\text{O}$$

$$-818.2 \text{ kJ} = q_{\text{H}_2\text{O}} + q_{\text{N}_2}$$

$$q_{\text{H}_2\text{O}} = 4 \text{ mol} \cdot 33.6 \text{ J} \cdot \Delta T$$

$$q_{\text{N}_2} = 1 \text{ mol} \cdot 29.1 \text{ J} \cdot \Delta T$$

$$818.2 \cdot 10^3 \text{ J} = (4 \text{ mol} \cdot 33.6 \cdot \Delta T) + (29.1 \cdot \Delta T)$$

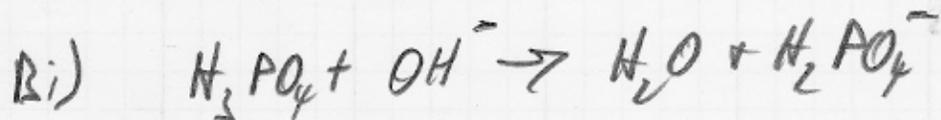
$$818.2 \cdot 10^3 \text{ J} = 163.5 \Delta T$$

$$\Delta T = 5.00 \cdot 10^3 \text{ }^\circ\text{C}$$

More Ch. 6 APQs



A ii) 2 moles of CO_2 for each C_2H_4

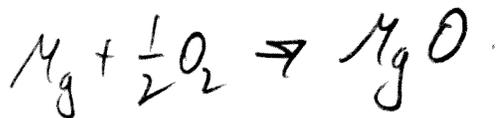


B ii) Basic $H_2PO_4^-$ would act as a Base



C ii) Green

2. Coffee cup calorimetry experiments can be used to obtain ΔH_f° for magnesium oxide.
 a. Write a balanced equation for the formation of magnesium oxide, whose enthalpy change is ΔH_f° .



b. To determine the heat capacity of the calorimeter, 49.6 mL of 1.01 M HCl are reacted with 50.1 mL of 0.998 M NaOH. The solution's temperature increases by 6.40°C. Determine the heat capacity of the calorimeter. You may assume the solution's specific heat capacity is 4.025 J·g⁻¹·°C⁻¹, the enthalpy of neutralization is -55.9 kJ per mole of H₂O and the solutions have a density of 1.0 g/mL



$$q_{\text{rxn}} = \frac{55.9 \text{ kJ}}{\text{mol}} \cdot 0.0500 \text{ mol} = -2.795 \cdot 10^3 \text{ J}$$

$$q_{\text{H}_2\text{O}} = 99.7 \text{ g} \cdot 4.025 \cdot 6.40^\circ\text{C} = 2.568 \cdot 10^3 \text{ J}$$

$$q_{\text{cal}} = 2.795 \cdot 10^3 - 2.568 \cdot 10^3 = 226.7 \text{ J}$$

$$C_{\text{cal}} = \frac{226.7}{6.40} = 35.4 \frac{\text{J}}{^\circ\text{C}}$$

c. When 0.221 g of magnesium turnings are added to 49.9 mL of 1.01 M HCl and 49.7 mL of H₂O in the same calorimeter, the temperature increases by 9.67°C. Write a balanced equation for the reaction that occurs and calculate the ΔH per mole magnesium. (Assume the solution's specific heat capacity is 3.862 J·g⁻¹·°C⁻¹ and the heat capacity of the calorimeter is the value obtained in b.).

$$\begin{array}{l}
 \overset{0.00909}{\text{Mg}} + \overset{0.050}{2\text{HCl}} \rightarrow \text{MgCl}_2 + \text{H}_2 \\
 -q_{\text{rxn}} = q_{\text{H}_2\text{O}} + q_{\text{cal}} = 2205 \\
 q_{\text{H}_2\text{O}} = 49.9\text{g} \cdot 3.862 \cdot 9.67 = 1865.5\text{J} \\
 50.12 \cdot 3.862 \cdot 9.67 = 1871.8\text{J} \\
 q_{\text{cal}} = 35.45 \cdot 9.67^\circ\text{C} = 342.3\text{J} \\
 \Delta H = \frac{4060\text{J} - 2205}{0.00909} = \frac{1855\text{J}}{0.00909} = 204000\text{J/mol} \\
 \Delta H = \frac{-2212}{0.00909} = -243.5\text{kJ/mol} \\
 -447\text{kJ/mol}
 \end{array}$$

d. When 0.576 g of MgO react with 51.0 mL of 1.01 M HCl and 50.1 mL of H₂O in the same calorimeter the temperature rises 4.72°C. Write a balanced equation for this reaction and calculate its ΔH per mole of MgO using the same assumptions as in part c.

$$\begin{array}{l}
 \overset{0.019\text{mol}}{\text{MgO}} + \overset{0.05\text{mol}}{2\text{HCl}} \rightarrow \text{MgCl}_2 + \text{H}_2\text{O} \\
 q_{\text{H}_2\text{O}} = 101.1\text{g} \cdot 3.862 \cdot 4.72 = 1853.5\text{J} \\
 q_{\text{cal}} = 35.4 \cdot 4.72 = 167.1\text{J} \\
 \Delta H = \frac{1853.5 - 2020}{0.019\text{mol}} = \frac{-166.5}{0.019} = -8763\text{J/mol} \\
 -144\text{kJ/mol}
 \end{array}$$

e. Use the above results and ΔH_f[°] of H₂O(l) (-285.8 kJ·mol⁻¹) to calculate ΔH_f[°] of magnesium oxide.

$$\begin{array}{l}
 \text{Mg} + \frac{1}{2}\text{O}_2 \rightarrow \text{MgO} \\
 \text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2 \quad \Delta H = -447\text{kJ/mol} \\
 \text{MgCl}_2 + \text{H}_2\text{O} \rightarrow \text{MgO} + 2\text{HCl} \quad \Delta H = -(-144\text{kJ/mol}) \\
 \text{H}_2 + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O} \quad \Delta H = -285.8\text{kJ/mol} \\
 \Delta H = -447 + 144 - 285.8 \\
 \Delta H = -588.8\text{kJ/mol}
 \end{array}$$