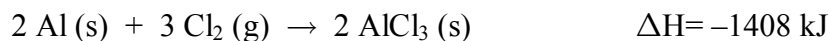


REVIEW QUESTIONS

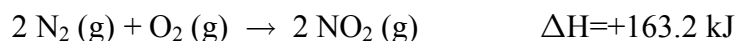
Chapter 6

1. How many kJ of heat is produced when 5.00 g of Al reacts with excess chlorine, as shown below:



$$5.00 \text{ g Al} \times \frac{1 \text{ mol}}{27.0 \text{ g}} \times \frac{1408 \text{ kJ}}{2 \text{ mol Al}} = 130. \text{ kJ}$$

2. Consider the following reaction:



- a) How many grams of N₂ must react to produce an enthalpy change of 5.00 kJ?

$$5.00 \text{ kJ} \times \frac{2 \text{ mol N}_2}{163.2 \text{ kJ}} \times \frac{28.0 \text{ g}}{1 \text{ mol}} = 1.72 \text{ g N}_2$$

- b) How much heat is transferred when 25.0 g of NO₂ is produced in this reaction?

$$25.0 \text{ g NO}_2 \times \frac{1 \text{ mol}}{46.0 \text{ g}} \times \frac{163.2 \text{ kJ}}{2 \text{ mol NO}_2} = 44.3 \text{ kJ}$$

3. When 6.50 g of solid NaOH is dissolved in 100.0 g of water in a calorimeter, the temperature rises from 21.6 °C to 37.8 °C. Calculate the change in enthalpy of solution in kJ/mol of NaOH. Assume specific heat of solution is the same as water.

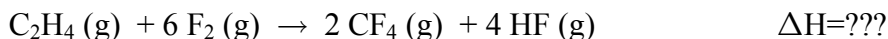
$$\text{mass of solution} = 100.0 \text{ g} + 6.50 \text{ g} = 106.5 \text{ g}$$

$$\Delta T = 37.8 - 21.6 = 16.2 \text{ }^\circ\text{C}$$

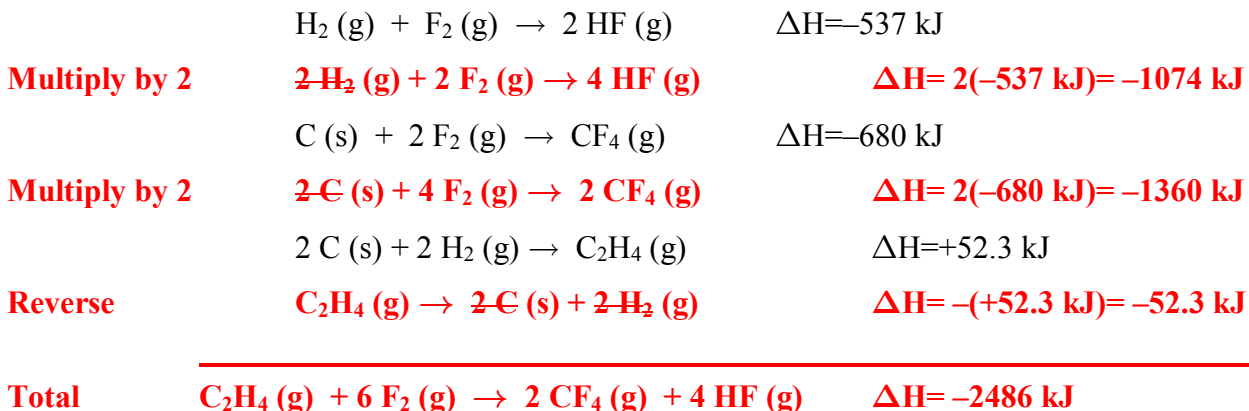
$$q = -(m \times s \times \Delta T) = -(106.5 \text{ g})(4.184 \text{ J/g }^\circ\text{C})(16.2 \text{ }^\circ\text{C}) = -7219 \text{ J}$$

$$\Delta H = \frac{-7219 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}}}{6.50 \text{ g NaOH} \times \frac{1 \text{ mol}}{40.0 \text{ g}}} = -44.4 \text{ kJ/mol NaOH}$$

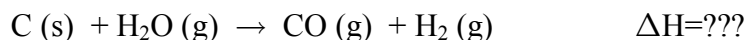
4. Calculate the enthalpy change for the reaction below:



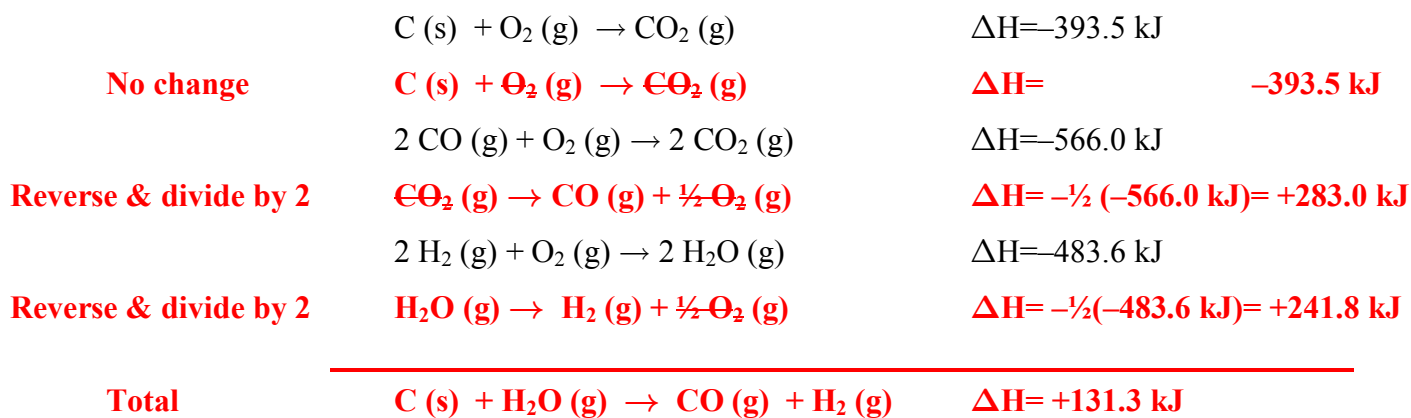
based on the following reactions:



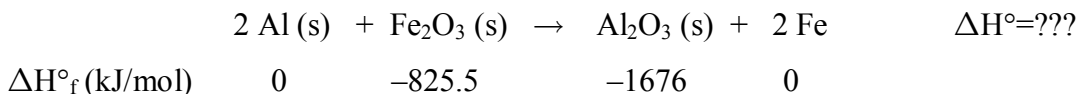
5. Determine the enthalpy change for the reaction below:



based on the following reactions:



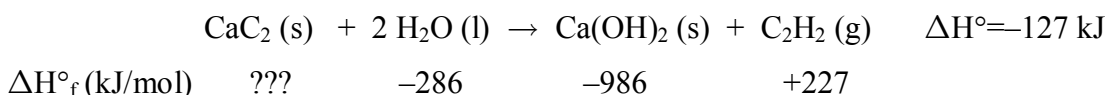
6. Use enthalpies of formation given to determine the standard enthalpy of reaction for the following:



$$\Delta H^\circ = \Sigma \Delta H^\circ \text{ (products)} - \Sigma \Delta H^\circ \text{ (reactants)}$$

$$\Delta H^\circ = [(-1676 \text{ kJ}) + 2(0)] - [2(0) + (-825.5 \text{ kJ})] = -851 \text{ kJ}$$

7. Use the enthalpies of formation and the enthalpy of reaction given below to determine the enthalpy of formation for solid CaC_2 .



$$\Delta H^\circ = \Sigma \Delta H^\circ \text{ (products)} - \Sigma \Delta H^\circ \text{ (reactants)}$$

$$\Delta H^\circ = [(-986 \text{ kJ}) + (+227 \text{ kJ})] - [(\text{???}) + 2(-286 \text{ kJ})] = -127 \text{ kJ}$$

$$\Delta H^\circ = [-759 \text{ kJ}] - [(\text{???}) + (-572)] = -127 \text{ kJ}$$

$$\text{???} = \Delta H^\circ_f (\text{CaC}_2) = -60 \text{ kJ/mol}$$

8. The heat of vaporization of a liquid (ΔH_{vap}) is the heat required to vaporize 1.00 g of a liquid. In one experiment, 60.0 g of liquid nitrogen (b.p. -196°C) are poured into a Styrofoam cup containing 200. g of water at 55.3°C . Calculate the molar heat of vaporization of liquid nitrogen if the final temperature of the water is 41.0°C .

$$\text{Heat lost by water} = \text{Heat gained by liquid nitrogen}$$

$$q_{\text{H}_2\text{O}} = m \times s \times \Delta T = (200. \text{ g})(4.184 \text{ J/g } ^\circ\text{C})(55.3^\circ\text{C} - 41.0^\circ\text{C}) = 11966 \text{ J}$$

$$\Delta H_{\text{vap}} = 11966 \text{ J} \times \frac{1 \text{ kJ}}{10^3 \text{ J}} \times \frac{1}{60.0 \text{ g}} \times \frac{28.0 \text{ g}}{1 \text{ mol}} = 5.58 \text{ kJ/mol}$$

9. Ice at 0 °C is placed in a Styrofoam cup containing 361 g of a soft drink at 23.0 °C. Some ice remains after the ice and the soft drink reach an equilibrium temperature of 0 °C. Determine the mass of the ice that has melted. Assume the specific heat of the drink to be the same as pure water. It requires 334 J of heat to melt 1 g of ice at 0 °C.

Heat lost by drink = Heat gained by ice

$$q_{\text{drink}} = m \times s \times \Delta T = (361 \text{ g})(4.184 \text{ J/g } ^\circ\text{C})(23.0 \text{ } ^\circ\text{C} - 0 \text{ } ^\circ\text{C}) = 34740 \text{ J}$$

$$\text{mass of ice melted} = 34740 \text{ J} \times \frac{1 \text{ g ice}}{334 \text{ J}} = 104 \text{ g}$$

10. A quantity of 200. mL of 0.862 M HCl is mixed with 200. mL of 0.431 M Ba(OH)₂ in a Styrofoam cup. The initial temperature of both solutions is 20.48 °C. Calculate the heat of neutralization in kJ/mol if the final temperature of the solution is 26.30 °C. Assume specific heat and density of solution to be the same as water.



$$\text{mass of solution} = 400. \text{ mL} \times \frac{1.00 \text{ g}}{1 \text{ mL}} = 400. \text{ g}$$

$$\Delta T = 26.30 \text{ } ^\circ\text{C} - 20.48 \text{ } ^\circ\text{C} = 5.82 \text{ } ^\circ\text{C}$$

$$q = -(m \times s \times \Delta T) = -(400. \text{ g})(4.184 \text{ J/g } ^\circ\text{C})(5.82 \text{ } ^\circ\text{C}) = -9740 \text{ J}$$

$$\text{mol H}_2\text{O} = 200. \text{ mL HCl} \times \frac{0.862 \text{ mol}}{1000 \text{ mL}} \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol HCl}} = 0.1724 \text{ mol}$$

$$\Delta H_{\text{neut}} = -9740 \text{ J} \times \frac{1 \text{ kJ}}{10^3 \text{ J}} \times \frac{1}{0.1724 \text{ mol H}_2\text{O}} = -56.5 \text{ kJ/mol}$$