

## Thermochemistry - Answers

1. Given:  $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightarrow 2 \text{NH}_3(\text{g}) + 92.5 \text{ kJ}$ , what is  $\Delta H_f^\circ$  for  $\text{NH}_3(\text{g})$ ?

$$\Delta H_{\text{rxn}} = (2 \text{ mol})(\text{NH}_3) - (1 \text{ mol})(\text{N}_2) - (3 \text{ mol})(\text{H}_2)$$

$$-92.5 \text{ kJ} = (2 \text{ mol})(\text{NH}_3) - (1 \text{ mol})(0 \text{ kJ/mol}) - (3 \text{ mol})(0 \text{ kJ/mol})$$

$$-92.5 \text{ kJ} = (2 \text{ mol})(\text{NH}_3)$$

$$\Delta H_f \text{ for } \text{NH}_3 = -46.2 \text{ kJ/mol}$$

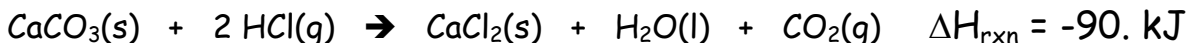
2. Given:
- |  |                              |
|--|------------------------------|
| $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$   | $\Delta H = 175 \text{ kJ}$  |
| $\text{Ca}(\text{OH})_2(\text{s}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CaO}(\text{s})$                               | $\Delta H = 67 \text{ kJ}$   |
| $\text{Ca}(\text{OH})_2(\text{s}) + 2 \text{HCl}(\text{g}) \rightarrow \text{CaCl}_2(\text{s}) + 2 \text{H}_2\text{O}(\text{l})$ | $\Delta H = -198 \text{ kJ}$ |

Calculate  $\Delta H_{\text{rxn}}$  for  $\text{CaCO}_3(\text{s}) + 2 \text{HCl}(\text{g}) \rightarrow \text{CaCl}_2(\text{s}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$

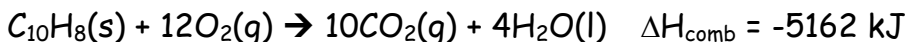
$$\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g}) \quad \Delta H = (1)(175 \text{ kJ})$$

$$\text{Ca}(\text{OH})_2(\text{s}) + 2 \text{HCl}(\text{g}) \rightarrow \text{CaCl}_2(\text{s}) + 2 \text{H}_2\text{O}(\text{l}) \quad \Delta H = (1)(-198 \text{ kJ})$$

$$\text{H}_2\text{O}(\text{l}) + \text{CaO}(\text{s}) \rightarrow \text{Ca}(\text{OH})_2(\text{s}) \quad \Delta H = (-1)(67 \text{ kJ})$$



3. The heat of combustion of naphthalene,  $\text{C}_{10}\text{H}_8(\text{s})$  is  $-5162 \text{ kJ}$ . What is the heat of formation of naphthalene?



$$\Delta H_{\text{comb}} = (10 \text{ mol})(\text{CO}_2) + (4 \text{ mol})(\text{H}_2\text{O}) - (1 \text{ mol})(\text{C}_{10}\text{H}_8) - (12 \text{ mol})(\text{O}_2)$$

$$-5162 \text{ kJ} = (10 \text{ mol})(-393.5 \text{ kJ/mol}) + (4 \text{ mol})(-285.8 \text{ kJ/mol}) - (1 \text{ mol})(\text{C}_{10}\text{H}_8) - (12 \text{ mol})(0 \text{ kJ/mol})$$

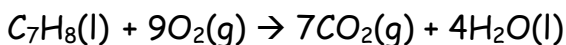
$$-5162 \text{ kJ} = -3935.0 \text{ kJ} - 1143.2 \text{ kJ} - (1 \text{ mol})(\text{C}_{10}\text{H}_8) - 0 \text{ kJ}$$

$$-5162 \text{ kJ} + 1143.2 \text{ kJ} + 3935.0 \text{ kJ} = -(1 \text{ mol})(\text{C}_{10}\text{H}_8)$$

$$-83.8 \text{ kJ} = -(1 \text{ mol})(\text{C}_{10}\text{H}_8)$$

$$\Delta H_f \text{ for } \text{C}_{10}\text{H}_8 = +83.8 \text{ kJ/mol}$$

4. Calculate the amount of heat given off when 100.0 g of  $C_7H_8(l)$  burns. The:  $\Delta H_f = 50.0 \text{ kJ/mol}$  for  $C_7H_8(l)$ ,



$$\Delta H_{\text{comb}} = (7 \text{ mol})(CO_2) + (4 \text{ mol})(H_2O) - (1 \text{ mol})(C_7H_8) - (9 \text{ mol})(O_2)$$

$$\Delta H_{\text{comb}} = (7 \text{ mol})(-393.5 \text{ kJ/mol}) + (4 \text{ mol})(-285.8 \text{ kJ/mol}) - (1 \text{ mol})(50.0 \text{ kJ/mol}) - (9 \text{ mol})(0 \text{ kJ/mol})$$

$$\Delta H_{\text{comb}} = -2754.5 \text{ kJ} - 1143.2 \text{ kJ} - 50.0 \text{ kJ}$$

$$\Delta H_{\text{comb}} = -3947.7 \text{ kJ}$$

$$\text{mol} = \frac{100.0 \text{ g}}{92.14052 \text{ g/mol}} = 1.085 \text{ mol}$$

$$\frac{1 \text{ mol}}{1.085 \text{ mol}} = \frac{-3947.7 \text{ kJ}}{x}$$

$$x = -4280 \text{ kJ for } 100.0 \text{ g of } C_7H_8$$

5. The standard enthalpy of combustion of  $C_2H_2(g)$  is  $-1301 \text{ kJ/mol}$ . What is the standard enthalpy of formation for  $C_2H_2(g)$ ?



$$\Delta H_{\text{comb}} = (4 \text{ mol})(CO_2) + (2 \text{ mol})(H_2O) - (2 \text{ mol})(C_2H_2) - (5 \text{ mol})(O_2)$$

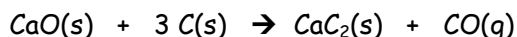
$$-2602 \text{ kJ} = (4 \text{ mol})(-393.5 \text{ kJ/mol}) + (2 \text{ mol})(-285.8 \text{ kJ/mol}) - (2 \text{ mol})(C_2H_2) - (5 \text{ mol})(0 \text{ kJ/mol})$$

$$-2602 \text{ kJ} = -1574.0 \text{ kJ} - 571.6 \text{ kJ} - (2 \text{ mol})(C_2H_2) - 0 \text{ kJ}$$

$$-456.4 \text{ kJ} = - (2 \text{ mol})(C_2H_2)$$

$$\Delta H_f C_2H_2 = +228 \text{ kJ/mol}$$

6. How much heat will be required to make 1.00 kg of  $\text{CaC}_2(\text{s})$  according to the reaction given below? The  $\Delta H_f$  value for  $\text{CaC}_2(\text{s})$  is  $-63 \text{ kJ/mol}$ .



$$\Delta H_{\text{rxn}} = (1 \text{ mol})(\text{CaC}_2) + (1 \text{ mol})(\text{CO}) - (1 \text{ mol})(\text{CaO}) - (3 \text{ mol})(\text{C})$$

$$\Delta H_{\text{rxn}} = (1 \text{ mol})(-63 \text{ kJ/mol}) + (1 \text{ mol})(-110.5 \text{ kJ/mol}) - (1 \text{ mol})(-635.1 \text{ kJ/mol}) - (3 \text{ mol})(0 \text{ kJ/mol})$$

$$\Delta H_{\text{rxn}} = -63 \text{ kJ} - 110.5 \text{ kJ} + 635.1 \text{ kJ} - 0 \text{ kJ}$$

$$\Delta H_{\text{rxn}} = +461.6 \text{ kJ}$$

$$\text{mol} = \frac{1000.0 \text{ g}}{64.100 \text{ g/mol}} = 15.6 \text{ mol}$$

$$\frac{1 \text{ mol}}{15.6 \text{ mol}} = \frac{461.6 \text{ kJ}}{x}$$

$$x = 7200 \text{ kJ for 1.00 kg of CaC}_2$$

7. 150.0 mL of 0.200 mol/L HCl was added to 150.0 mL of 0.200 mol/L NaOH in a styrene cup. Initial temperature of both solutions was  $25.0^\circ\text{C}$ . The final temperature was  $27.1^\circ\text{C}$ . Calculate the heat of reaction per mole of hydrochloric acid.

Heat released by the reaction = heat absorbed by calorimeter and water

$$\Delta H = mc\Delta T (\text{calorimeter}) + mc\Delta T (\text{water} - \text{HCl}(\text{aq})) + mc\Delta T (\text{water} - \text{NaOH}(\text{aq}))$$

assume the heat absorbed by the coffee cup calorimeter is negligible

$$\Delta H = mc\Delta T (\text{water} - \text{HCl}(\text{aq})) + mc\Delta T (\text{water} - \text{NaOH}(\text{aq}))$$

$$\Delta H = (150.0 \text{ g})(4.18 \text{ g/J}^\circ\text{C})(2.1^\circ\text{C}) + (150.0 \text{ g})(4.18 \text{ g/J}^\circ\text{C})(2.1^\circ\text{C})$$

$$\Delta H = 1316.7 \text{ J} + 1316.7 \text{ J}$$

$$\Delta H = 2633.4 \text{ J absorbed by the calorimeter and water}$$

$$\Delta H = -2633.4 \text{ J for 50.0 mL of 1.00 mol/L HCl}$$

$$n = CV = (0.200 \text{ mol/L})(0.1500 \text{ L}) = 0.0300 \text{ mol HCl}$$

$$\frac{0.0300 \text{ mol}}{1.00 \text{ mol}} = \frac{-2.6334 \text{ kJ}}{x}$$

$$x = -88 \text{ kJ/mol} = \Delta H_{\text{rxn}}$$

8. In an aluminum calorimeter, 20.0 g of nitrogen was burned in oxygen to produce nitrogen monoxide. From the following data, find the heat of reaction per mole of nitrogen burned.
- Mass of nitrogen burned                      20.0 g
  - Mass of aluminum calorimeter              70.37 g
  - Volume of water in calorimeter            500.0 mL
  - Initial temperature                            72.60°C
  - Final temperature                             82.80°C

Heat released by the reaction = heat absorbed by calorimeter and water

$$\Delta H = mc\Delta T (\text{calorimeter}) + mc\Delta T (\text{water})$$

$$\Delta H = (0.07037 \text{ kg})(0.900 \text{ kJ/kg}^\circ\text{C})(10.2^\circ\text{C}) + (0.500 \text{ kg})(4.18 \text{ kJ/kg}^\circ\text{C})(10.2^\circ\text{C})$$

$$\Delta H = 0.646 \text{ kJ} + 21.3 \text{ kJ}$$

$$\Delta H = 21.94 \text{ kJ absorbed by the calorimeter and water}$$

$$\Delta H = -21.94 \text{ kJ for 20.0 g of nitrogen}$$

$$\text{mol} = \frac{20.0 \text{ g}}{28.0134 \text{ g/mol}} = 0.7139 \text{ mol}$$

$$\frac{0.7139 \text{ mol}}{1.00 \text{ mol}} = \frac{-21.94 \text{ kJ}}{x}$$

$$x = -30.8 \text{ kJ/mol} = \Delta H_{\text{rxn}}$$

9. In a styrene cup, some solid copper was oxidized in excess oxygen to produce copper(II) oxide. Use the data below to establish the heat of reaction per mole of copper.
- Mass of copper 1.00 g
  - Mass of water in calorimeter 100.0 g
  - Initial temperature of water 21.0°C
  - Final temperature of water 26.9°C

Heat released by the reaction = heat absorbed by calorimeter and water

$$\Delta H = mc\Delta T (\text{calorimeter}) + mc\Delta T (\text{water})$$

assume the amount of heat absorbed by the styrene cup is negligible

$$\Delta H = mc\Delta T (\text{water})$$

$$\Delta H = (0.100 \text{ kg})(4.18 \text{ kJ/kg}^\circ\text{C})(5.9^\circ\text{C})$$

$$\Delta H = 2.466 \text{ kJ absorbed by the water}$$

$$\Delta H = -2.466 \text{ kJ for 1.00 g of copper}$$

$$\text{mol} = \frac{1.00 \text{ g}}{63.546 \text{ g/mol}} = 0.0157 \text{ mol}$$

$$\frac{0.0157 \text{ mol}}{1.00 \text{ mol}} = \frac{-2.466 \text{ kJ}}{x}$$

$$x = -160 \text{ kJ/mol} = \Delta H_{\text{rxn}}$$

10. In an aluminum calorimeter, 25.00 g of iodine was burned in excess hydrogen to produce hydrogen iodide gas. From the data below, establish the heat of reaction per mole of iodine.
- Mass of iodine 25.00 g
  - Mass of aluminum calorimeter 70.0 g
  - Volume of water in calorimeter 100.0 mL
  - Initial temperature 23.8°C
  - Final temperature 34.6°C

Heat released by the reaction = heat absorbed by calorimeter and water

$$\Delta H = mc\Delta T (\text{calorimeter}) + mc\Delta T (\text{water})$$

$$\Delta H = (0.0700 \text{ kg})(0.900 \text{ kJ/kg}^\circ\text{C})(10.8^\circ\text{C}) + (0.100 \text{ kg})(4.18 \text{ kJ/kg}^\circ\text{C})(10.8^\circ\text{C})$$

$$\Delta H = 0.6804 \text{ kJ} + 4.5144 \text{ kJ}$$

$$\Delta H = 5.1948 \text{ kJ absorbed by the calorimeter and water}$$

$$\Delta H = -5.1948 \text{ kJ for 25.00 g of iodine}$$

$$\text{mol} = \frac{25.00 \text{ g}}{253.809 \text{ g/mol}} = 0.098499 \text{ mol}$$

$$\frac{0.098499 \text{ mol}}{1.00 \text{ mol}} = \frac{-5.1948 \text{ kJ}}{x}$$

$$x = -52.7 \text{ kJ/mol} = \Delta H_{\text{rxn}}$$

11. Sodium was burned in oxygen to burn sodium oxide. An aluminum calorimeter was used. Calculate the heat of reaction per mole of sodium oxide formed.

▪ Mass of sodium	1.00 g
▪ Mass of aluminum calorimeter	25.31 g
▪ Volume of water in calorimeter	350.0 mL
▪ Initial temperature	27.51°C
▪ Final temperature	31.82°C

Heat released by the reaction = heat absorbed by calorimeter and water

$$\Delta H = mc\Delta T (\text{calorimeter}) + mc\Delta T (\text{water})$$

$$\Delta H = (0.02531 \text{ kg})(0.900 \text{ kJ/kg}^\circ\text{C})(4.31^\circ\text{C}) + (0.350 \text{ kg})(4.18 \text{ kJ/kg}^\circ\text{C})(4.31^\circ\text{C})$$

$$\Delta H = 0.09818 \text{ kJ} + 6.306 \text{ kJ}$$

$$\Delta H = 6.404 \text{ kJ absorbed by the calorimeter and water}$$

$$\Delta H = -6.404 \text{ kJ for 1.00 g of sodium}$$

$$\text{mol} = \frac{1.00 \text{ g}}{22.98977 \text{ g/mol}} = 0.043498 \text{ mol}$$

$$\frac{0.043498 \text{ mol}}{1.00 \text{ mol}} = \frac{-6.404 \text{ kJ}}{x}$$

$$x = -147 \text{ kJ/mol} = \Delta H_{\text{rxn}}$$

12. 0.844 g of formaldehyde (methanal),  $\text{HCHO}$ , reacted with oxygen in a Styrofoam calorimeter to produce carbon dioxide and water. The mass of water in the calorimeter was 150.0 g and the temperature increased from  $21.8^\circ\text{C}$  to  $45.1^\circ\text{C}$ . Calculate the heat of reaction per mole of formaldehyde burned.

Heat released by the reaction = heat absorbed by calorimeter and water

$$\Delta H = mc\Delta T (\text{calorimeter}) + mc\Delta T (\text{water})$$

assume the amount of heat absorbed by the Styrofoam calorimeter is negligible

$$\Delta H = mc\Delta T (\text{water})$$

$$\Delta H = (0.150 \text{ kg})(4.18 \text{ kJ/kg}^\circ\text{C})(23.3^\circ\text{C})$$

$$\Delta H = 14.61 \text{ kJ absorbed by the water}$$

$$\Delta H = -14.61 \text{ kJ for 0.844 g of formaldehyde}$$

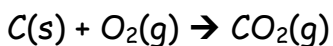
$$\text{mol} = \frac{0.844 \text{ g}}{30.02628 \text{ g/mol}} = 0.0281 \text{ mol}$$

$$\frac{0.0281 \text{ mol}}{1.00 \text{ mol}} = \frac{-14.61 \text{ kJ}}{x}$$

$$x = -519 \text{ kJ/mol} = \Delta H_{\text{comb}} \text{ for formaldehyde}$$



13. Exactly 3.00 g of  $C(s)$  was burned to  $CO_2(g)$  in a copper calorimeter. The mass of the calorimeter was 1.500 kg and the mass of the water in which the calorimeter was immersed was 2.000 kg. The initial temperature of the system was  $20.0^\circ C$  and the final temperature was  $31.0^\circ C$ . Calculate the heat of formation of  $CO_2(g)$ , under the conditions present in the calorimeter.



Heat released by the reaction = heat absorbed by calorimeter and water

$$\Delta H = mc\Delta T (\text{calorimeter}) + mc\Delta T (\text{water})$$

$$\Delta H = (1.500 \text{ kg})(0.385 \text{ kJ/kg}^\circ C)(11.0^\circ C) + (2.000 \text{ kg})(4.18 \text{ kJ/kg}^\circ C)(11.0^\circ C)$$

$$\Delta H = 6.33525 \text{ kJ} + 92.048 \text{ kJ}$$

$$\Delta H = 98.41 \text{ kJ absorbed by calorimeter and water}$$

$\therefore$  98.41 kJ were released by the combustion of 3.00 g of carbon

$$\text{mol} = \text{mass} \div \text{molar mass}$$

$$\text{mol} = 3.00 \text{ g} \div 12.011 \text{ g/mol} = 0.2498 \text{ mol } C$$

$$= 0.2498 \text{ mol } CO_2$$

$$\frac{-98.41 \text{ kJ}}{x} = \frac{0.2498 \text{ mol } CO_2}{1 \text{ mol}}$$

$$\Delta H_f \text{ for } CO_2 = -394 \text{ kJ/mol}$$

$$x = -394 \text{ kJ/mol}$$

14. The specific heat capacity of Ni(s) is 0.444 kJ/Kg°C. A 3.85 g sample of benzoic acid,  $C_6H_5COOH(s)$ , was burned in a nickel calorimeter having a mass of 0.850 kg and immersed in 1.200 kg of water. The initial temperature of the system was 23.0°C and the final temperature was 41.0°C. Calculate the heat of combustion of benzoic acid, under the conditions present inside the calorimeter.

Heat released by the reaction = heat absorbed by calorimeter and water

$$\Delta H = mc\Delta T (\text{calorimeter}) + mc\Delta T (\text{water})$$

$$\Delta H = (0.850 \text{ kg})(0.444 \text{ kJ/kg}^\circ\text{C})(18.0^\circ\text{C}) + (1.200 \text{ kg})(4.18 \text{ kJ/kg}^\circ\text{C})(18.0^\circ\text{C})$$

$$\Delta H = 6.7932 \text{ kJ} + 90.288 \text{ kJ}$$

$$\Delta H = 97.0812 \text{ kJ absorbed by the calorimeter and water}$$

$$\Delta H = -97.0812 \text{ kJ for 3.85 g of benzoic acid}$$

$$\text{mol} = \frac{3.85 \text{ g}}{122.12344 \text{ g/mol}} = 0.0315 \text{ mol}$$

$$\frac{0.0315 \text{ mol}}{1.00 \text{ mol}} = \frac{-97.0812 \text{ kJ}}{x}$$

$$x = -3080 \text{ kJ/mol} = \Delta H_{\text{comb}} \text{ for benzoic acid}$$

15. A 2.50g sample of sucrose,  $C_{12}H_{22}O_{11}(s)$  was burned in a 2.100 kg iron calorimeter immersed in 1.450 kg of water. The initial temperature of the system was 24.32°C and the final temperature was 30.20°C. Determine the heat of combustion of sucrose, under the conditions present inside the calorimeter.

Heat released by the reaction = heat absorbed by calorimeter and water

$$\Delta H = mc\Delta T (\text{calorimeter}) + mc\Delta T (\text{water})$$

$$\Delta H = (2.100 \text{ kg})(0.444 \text{ kJ/kg}^\circ\text{C})(5.88^\circ\text{C}) + (1.450 \text{ kg})(4.18 \text{ kJ/kg}^\circ\text{C})(5.88^\circ\text{C})$$

$$\Delta H = 5.482512 \text{ kJ} + 35.63868 \text{ kJ}$$

$$\Delta H = 41.12 \text{ kJ absorbed by the calorimeter and water}$$

$$\Delta H = -41.121 \text{ kJ for 2.50 g of sucrose}$$

$$\text{mol} = \frac{2.50 \text{ g}}{342.30008 \text{ g/mol}} = 0.00730 \text{ mol}$$

$$\frac{0.00730 \text{ mol}}{1.00 \text{ mol}} = \frac{-41.12 \text{ kJ}}{x}$$

$$x = -5630 \text{ kJ/mol} = \Delta H_{\text{comb}} \text{ for sucrose}$$

16. 50.0 mL of 0.800 mol/L hydrobromic acid was added to 50.0 mL of 0.800 mol/L potassium hydroxide in a styrene cup. Initial temperature of both solutions was 23.18°C. Final temperature was 26.38°C. Calculate the heat of reaction per mole of hydrobromic acid.

Heat released by the reaction = heat absorbed by calorimeter and water

$$\Delta H = mc\Delta T (\text{calorimeter}) + mc\Delta T (\text{water} - \text{HBr(aq)}) + mc\Delta T (\text{water} - \text{KOH(aq)})$$

assume the heat absorbed by the coffee cup calorimeter is negligible

$$\Delta H = mc\Delta T (\text{water} - \text{HBr(aq)}) + mc\Delta T (\text{water} - \text{KOH(aq)})$$

$$\Delta H = (50.0 \text{ g})(4.18 \text{ g/J}^\circ\text{C})(3.20^\circ\text{C}) + (50.0 \text{ g})(4.18 \text{ g/J}^\circ\text{C})(3.20^\circ\text{C})$$

$$\Delta H = 668.8 \text{ J} + 668.8 \text{ J}$$

$$\Delta H = 1337.6 \text{ J absorbed by the calorimeter and water}$$

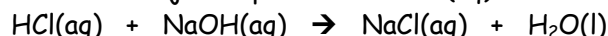
$$\Delta H = -1337.6 \text{ J for 50.0 mL of 1.00 mol/L HCl}$$

$$n = CV = (0.800 \text{ mol/L})(0.0500 \text{ L}) = 0.0400 \text{ mol HCl}$$

$$\frac{0.0400 \text{ mol}}{1.00 \text{ mol}} = \frac{-1.3376 \text{ kJ}}{x}$$

$$x = -33.4 \text{ kJ/mol} = \Delta H_{\text{rxn}}$$

17. When 50.0 mL of 1.00 mol/L HCl(aq) and 50.0 mL of 1.00 mol/L NaOH(aq) are mixed in a "coffee cup" calorimeter, the temperature of the resulting solution increases from 21.0°C to 27.5°C. Use the specific heat capacity of H<sub>2</sub>O(l), given on page one of the Chemistry Data Booklet, to calculate the heat of this reaction measured in kilojoules per mole of HCl(aq).



Heat released by the reaction = heat absorbed by calorimeter and water

$$\Delta H = mc\Delta T (\text{calorimeter}) + mc\Delta T (\text{water} - \text{HCl(aq)}) + mc\Delta T (\text{water} - \text{NaOH(aq)})$$

assume the heat absorbed by the coffee cup calorimeter is negligible

$$\Delta H = mc\Delta T (\text{water} - \text{HCl(aq)}) + mc\Delta T (\text{water} - \text{NaOH(aq)})$$

$$\Delta H = (50.0 \text{ g})(4.18 \text{ g/J}^\circ\text{C})(6.5^\circ\text{C}) + (50.0 \text{ g})(4.18 \text{ g/J}^\circ\text{C})(6.5^\circ\text{C})$$

$$\Delta H = 1358.5 \text{ J} + 1358.5 \text{ J}$$

$$\Delta H = 2717 \text{ J absorbed by the calorimeter and water}$$

$$\Delta H = -2717 \text{ J for 50.0 mL of 1.00 mol/L HCl}$$

$$n = CV = (1.00 \text{ mol/L})(0.0500 \text{ L}) = 0.0500 \text{ mol HCl}$$

$$\frac{0.0500 \text{ mol}}{1.00 \text{ mol}} = \frac{-2.717 \text{ kJ}}{x}$$

$$x = -54.3 \text{ kJ/mol} = \Delta H_{\text{rxn}}$$

18. We often have to make assumptions when we conduct laboratory experiments. Consider carefully the experiment described in question 17, and list four unstated assumptions that must be made in order for one to determine the heat of reaction.

1. We assume that the reaction happens fast enough that no heat is transferred from the system to the environment or vice versa.
2. We assume that the densities of HCl(aq), NaOH(aq), and NaCl(aq) all equal 1.00 g/mol.
3. We assume that it does take 4.184 J to increase the temperature of the solution by 1.00°C just as it does for pure water.
4. We assume that we can ignore the heat capacity of the material from which the "coffee cup" is made and that the total heat capacity of the calorimeter is just the heat capacity of the liquid contents.