Thermochemistry - Answers

1. Given: $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g) + 92.5\ KJ$, what is ΔH_f° for $NH_3(g)$?

$\Delta H_{rxn} = (2\ mol)(NH_3) - (1\ mol)(N_2) - (3\ mol)(H_2)$

$-92.5\ kJ = (2\ mol)(NH_3) - (1\ mol)(0\ kJ/mol) - (3\ mol)(0\ kJ/mol)$

$-92.5\ kJ = (2\ mol)(NH_3)$

ΔH_f for $NH_3 = -46.2\ kJ/mol$

2. Given:

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ $\Delta H = 175\ kJ$

$Ca(OH)_2(s) \rightarrow H_2O(l) + CaO(s)$ $\Delta H = 67\ kJ$

$Ca(OH)_2(s) + 2\ HCl(g) \rightarrow CaCl_2(s) + 2\ H_2O(l)$ $\Delta H = -198\ kJ$

Calculate ΔH_{rxn} for $CaCO_3(s) + 2\ HCl(g) \rightarrow CaCl_2(s) + H_2O(l) + CO_2(g)$

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ $\Delta H = (1)(175\ kJ)$

$Ca(OH)_2(s) + 2\ HCl(g) \rightarrow CaCl_2(s) + 2\ H_2O(l)$ $\Delta H = (1)(-198\ kJ)$

$\underline{H_2O(l) + CaO(s) \rightarrow Ca(OH)_2(s) \qquad \Delta H = (-1)(67\ kJ)}$

$CaCO_3(s) + 2\ HCl(g) \rightarrow CaCl_2(s) + H_2O(l) + CO_2(g)$ $\Delta H_{rxn} = -90.\ kJ$

3. The heat of combustion of naphthalene, $C_{10}H_8(s)$ is $-5162\ kJ$. What is the heat of formation of naphthalene?

$C_{10}H_8(s) + 12O_2(g) \rightarrow 10CO_2(g) + 4H_2O(l)$ $\Delta H_{comb} = -5162\ kJ$

$\Delta H_{comb} = (10\ mol)(CO_2) + (4\ mol)(H_2O) - (1\ mol)(C_{10}H_8) - (12\ mol)(O_2)$

$-5162\ kJ = (10\ mol)(-393.5\ kJ/mol) + (4\ mol)(-285.8\ kJ/mol) - (1\ mol)(C_{10}H_8) - (12\ mol)(0\ kJ/mol)$

$-5162\ kJ = -3935.0\ kJ - 1143.2\ kJ - (1\ mol)(C_{10}H_8) - 0\ kJ$

$-5162\ kJ + 1143.2\ kJ + 3935.0\ kJ = -(1\ mol)(C_{10}H_8)$

$-83.8\ kJ = -(1\ mol)(C_{10}H_8)$

ΔH_f for $C_{10}H_8 = +83.8\ kJ/mol$

4. Calculate the amount of heat given off when 100.0 g of $C_7H_8(l)$ burns. The: ΔH_f = 50.0 kJ/mol for $C_7H_8(l)$,

$$C_7H_8(l) + 9O_2(g) \rightarrow 7CO_2(g) + 4H_2O(l)$$

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

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

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

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

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

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

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

5. The standard enthalpy of combustion of $C_2H_2(g)$ is –1301 kJ/mol. What is the standard enthalpy of formation for $C_2H_2(g)$?

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$$

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

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

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

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

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

6. How much heat will be required to make 1.00 kg of $CaC_2(s)$ according to the reaction given below? The ΔH_f value for $CaC_2(s)$ is –63 kJ/mol.

$$CaO(s) + 3\,C(s) \rightarrow CaC_2(s) + CO(g)$$

ΔH_{rxn} = (1 mol)(CaC_2) + (1 mol)(CO) – (1 mol)(CaO) – (3 mol)(C)

ΔH_{rxn} = (1 mol)(-63 kJ/mol) + (1 mol)(-110.5 kJ/mol) – (1 mol)(-635.1 kJ/mol) – (3 mol)(0 kJ/mol)

ΔH_{rxn} = -63 kJ – 110.5 kJ + 631.5 kJ – 0 kJ

ΔH_{rxn} = +461.6 kJ

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

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

$$x = 7200\text{ kJ for }1.00\text{ kg of }CaC_2$$

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˚C. The final temperature was 27.1˚C. Calculate the heat of reaction per mole of hydrochloric acid.

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

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

assume the heat absorbed by the coffee cup calorimeter is negligible

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

ΔH = (150.0 g)(4.18 g/J°C)(2.1°C) + (150.0 g)(4.18 g/J°C)(2.1°C)

ΔH = 1316.7 J + 1316.7 J

ΔH = 2633.4 J absorbed by the calorimeter and water

ΔH = -2633.4 J for 50.0 mL of 1.00 mol/L HCl

n=CV = (0.200 mol/L)(0.1500 L) = 0.0300 mol HCl

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

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

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

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

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

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

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

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

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

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

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

9. In a styrene cup, come 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

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

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

$\Delta H = mc\Delta T$ (water)

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

$\Delta H = 2.466$ kJ absorbed by the water

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

12. 0.844 g of formaldehyde (methanal), 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°C to 45.1°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

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

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

$\Delta H = 14.61$ kJ absorbed by the water

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

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

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

$x = -519\text{ kJ/mol} = \Delta H_{comb}$ 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°C and the final temperature was 31.0°C. Calculate the heat of formation of $CO_2(g)$, under the conditions present in the calorimeter.

$C(s) + O_2(g) \rightarrow CO_2(g)$

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

$\Delta H = mc\Delta T$ (calorimeter) + $mc\Delta T$ (water)
$\Delta H = (1.500\ \text{kg})(0.385\ \text{kg/kJ°C})(11.0°C) + (2.000\ \text{kg})(4.18\ \text{kJ/kg°C})(11.0°C)$
$\Delta H = 6.33525\ \text{kJ} + 92.048\ \text{kJ}$
$\Delta H = 98.41$ kJ absorbed by calorimeter and water

∴ 98.41 kJ were released by the combustion of 3.00 *g* of carbon

mol = mass ÷ molar mass
mol = 3.00 *g* ÷ 12.011 *g*/mol = 0.2498 mol *C*
= 0.2498 mol CO_2

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

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

ΔH_f for CO_2 = -394 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

$\Delta H = mc\Delta T$ (calorimeter) + $mc\Delta T$ (water)
$\Delta H = (0.850\text{ kg})(0.444\text{ kg/kJ}^\circ C)(18.0^\circ C) + (1.200\text{ kg})(4.18\text{ kJ/kg}^\circ C)(18.0^\circ C)$
$\Delta H = 6.7932\text{ kJ} + 90.288\text{ kJ}$
$\Delta H = 97.0812$ kJ absorbed by the calorimeter and water
$\Delta H = -97.0812$ 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

$\Delta H = mc\Delta T$ (calorimeter) + $mc\Delta T$ (water)
$\Delta H = (2.100\text{ kg})(0.444\text{ kg/kJ}^\circ C)(5.88^\circ C) + (1.450\text{ kg})(4.18\text{ kJ/kg}^\circ C)(5.88^\circ C)$
$\Delta H = 5.482512\text{ kJ} + 35.63868\text{ kJ}$
$\Delta H = 41.12$ kJ absorbed by the calorimeter and water
$\Delta H = -41.121$ 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

$\Delta H = mc\Delta T$ (calorimeter) + $mc\Delta T$ (water - $HBr(aq)$) + $mc\Delta T$ (water - $KOH(aq)$)
assume the heat absorbed by the coffee cup calorimeter is negligible

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

$\Delta H = (50.0\ g)(4.18\ g/J^\circ C)(3.20^\circ C) + (50.0\ g)(4.18\ g/J^\circ C)(3.20^\circ C)$
$\Delta H = 668.8\ J + 668.8\ J$
$\Delta H = 1337.6\ J$ absorbed by the calorimeter and water
$\Delta H = -1337.6\ J$ for 50.0 mL of 1.00 mol/L HCl

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

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

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

17. When 50.0mL of 1.00mol/L HCl(aq) and 50.0 mL of 1.00mol/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 $H_2O(l)$, given on page one of the Chemistry Data Booklet, to calculate the heat of this reaction measured in kilojoules per mole of HCl(aq).

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

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

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

assume the heat absorbed by the coffee cup calorimeter is negligible

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

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

$\Delta H = 1358.5\ J + 1358.5\ J$

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

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

n=CV = (1.00 mol/L)(0.0500 L) = 0.0500 mol HCl

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

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

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.