Thermochemistry Review – Answers

1. Using the data below, calculate the amount of energy required to melt a 16.2 g aluminum pop can. Assume that the aluminum is at an initial temperature of 20°C.
   - Molar heat of fusion of Al(s) = 10.7 kJ/mol
   - Melting point of Al(s) = 660°C
   - Heat capacity of Al(s) = 0.900 J/g°C

$\text{mol} = 16.2\ \text{g} \div 26.98154\ \text{g/mol} = 0.600\ \text{mol}$

$\Delta H = (20°C \rightarrow 660°C) + (\text{melt})$

$\Delta H = (mc\Delta T) + (n\Delta H_{fus})$

$\Delta H = (16.2\ \text{g})(0.900\ \text{J/g°C})(640°C) + (0.600\ \text{mol})(10.7\ \text{kJ/mol})$

$\Delta H = 9331.2\ \text{J} + 6.42\ \text{kJ}$

$\Delta H = 9.33\ \text{kJ} + 6.42\ \text{kJ}$

$\boldsymbol{\Delta H = 15.8\ \text{kJ}}$

2. Given the thermochemical equations:

$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -394\ \text{kJ}$

$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l) \qquad \Delta H = -286\ \text{kJ}$

$C(s) + 2\ H_2(g) + \frac{1}{2} O_2(g) \rightarrow CH_3OH(l) \qquad \Delta H = -239\ \text{kJ}$

Calculate the ΔH for the reaction: $CH_3OH(l) + {}^{3}/_{2}\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$

$CH_3OH(l) \rightarrow C(s) + 2\ H_2(g) + \frac{1}{2} O_2(g) \qquad \Delta H = (-1)(-239\ \text{kJ})$

$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = (1)(-394\ \text{kJ})$

$\underline{2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l) \qquad \Delta H = (2)(-286\ \text{kJ})}$

$\boldsymbol{CH_3OH(l) + {}^{3}/_{2}\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l) \qquad \Delta H = -727\ \text{kJ}}$

3. Consider the two thermochemical equations:

$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l)$ $\Delta H = -286\text{ kJ}$
$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g)$ $\Delta H = -242\text{ kJ}$

a) Provide an explanation based on chemical bonding for the energy change that occurs in the ***first*** reaction.

In the first reaction, energy is required to break the chemical bonds in the hydrogen molecules and in the oxygen molecules. Energy is produced when hydrogen atoms bond to oxygen atoms as water molecules form. Since the reaction is exothermic, more energy is given off as H-O bonds form than is needed to break the H-H bonds and O-O bonds in the hydrogen and oxygen molecules.

b) Based on forces of attraction, explain the difference in ΔH values between these two reactions. (Value: 2)

The first reaction gives off more energy than the second because there is an additional exothermic process involved. As water molecules come together in the liquid phase, they are attracted to one another in clusters by relatively strong intermolecular forces of attraction. The difference in the ΔH values between these two reactions is the energy required to overcome the intermolecular forces of attraction as liquid water is transformed to gaseous water (steam).

4. Given the following thermochemical equations:

$PH_3(g) \rightarrow P(g) + 3H(g)$ $\Delta H = 965\text{ kJ}$
$O_2(g) \rightarrow 2\,O(g)$ $\Delta H = 490\text{ kJ}$
$2\,H(g) + O(g) \rightarrow H_2O(g)$ $\Delta H = -930\text{ kJ}$
$2\,P(g) + 5\,O(g) \rightarrow P_2O_5(g)$ $\Delta H = -3382\text{ kJ}$

calculate ΔH for the following reaction:
$2\,PH_3(g) + 4\,O_2(g) \rightarrow 3\,H_2O(g) + P_2O_5(g)$
(Value: 4)

$2PH_3(g) \rightarrow 2P(g) + 6H(g)$ $\Delta H = (2)(965\text{ kJ})$
$4O_2(g) \rightarrow 8O(g)$ $\Delta H = (4)(490\text{ kJ})$
$2P(g) + 5O(g) \rightarrow P_2O_5(g)$ $\Delta H = (1)(-3382\text{ kJ})$
$\underline{6H(g) + 3O(g) \rightarrow 3H_2O(g) \qquad \Delta H = (3)(-930\text{ kJ})}$

$\mathbf{2PH_3(g) + 4O_2(g) \rightarrow 3H_2O(g) + P_2O_5(g) \qquad \Delta H = -2280\text{ kJ}}$

5. Use these thermochemical equations to calculate the heat of formation of $C_2H_6(g)$

$$2\ C(s) + 3\ H_2(g) \rightarrow C_2H_6(g)$$

(Value: 4)

$C(s) + O_2(g) \rightarrow CO_2(g)$	$\Delta H = -394.8\ kJ$
$H_2(g) + \frac{1}{2}\ O_2(g) \rightarrow H_2O(l)$	$\Delta H = -286.1\ kJ$
$C_2H_6(g) + 3\frac{1}{2}\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(l)$	$\Delta H = -1559.3\ kJ$

$$2C(s) + 2O_2(g) \rightarrow 2CO_2(g) \qquad \Delta H = (2)(-394.8\ kJ)$$

$$3H_2(g) + {}^{3}/_{2}\ O_2(g) \rightarrow 3H_2O(l) \qquad \Delta H = (3)(-286.1\ kJ)$$

$$\underline{2\ CO_2(g) + 3\ H_2O(l) \rightarrow C_2H_6(g) + 3\tfrac{1}{2}\ O_2(g) \qquad \Delta H = (-1)(-1559.3\ kJ)}$$

$$\mathbf{2\ C(s) + 3\ H_2(g) \rightarrow C_2H_6(g) \qquad \Delta H = -88.6\ kJ}$$

6. Using the data below, calculate the total amount of energy required to change 25.0 g of solid mercury at –50.0°C to liquid mercury at 60.0°C.

(Value: 4)

- Molar heat of fusion of Hg(s) = 2.27 kJ/mol
- Melting point of Hg(s) = –38.8°C
- Specific heat capacity of Hg(s) = 0.141 J/(g°C)
- Specific heat capacity of Hg(l) = 0.138 J/(g°C)

$$\Delta H = (-50.0°C \rightarrow -38.8°C) + (\text{melt}) + (-38.8°C \rightarrow 60.0°C)$$

$$\Delta H = (mc\Delta T) + (n\Delta H_{fus}) + (mc\Delta T)$$

$$\Delta H = (0.0250\ kg)(0.141\ kJ/kg°C)(11.2°C) + (0.125\ mol)(2.27\ kJ/mol) + (0.0250\ kg)(0.138\ kJ/kg°C)(98.8°C)$$

$$\Delta H = 0.03948\ kJ + 0.283\ kJ + 0.34086\ kJ$$

$$\Delta H = 0.663255\ kJ$$

$$\Delta H = 0.663\ kJ \text{ or } 633\ J$$

7. The reaction that occurs when a typical fat, glycerol trioleate, is metabolized in the body is:

$$C_{57}H_{104}O_6(s) + 80\ O_2(g) \rightarrow 57\ CO_2(g) + 52\ H_2O(l)$$

a) 37.8 kJ is evolved when 1.00 g of this fat (molar mass = 884 g/mol) is metabolized. Determine the amount of energy evolved when one mole of this fat is metabolized.

$$\text{mol} = \frac{1.00\ g}{885.44916\ g/\text{mol}} = 0.00113\ \text{mol}$$

$$\frac{37.8\ kJ}{x} = \frac{0.00113\ \text{mol}}{1.00\ \text{mol}} \qquad x = 33.469.978$$

$$\text{energy evolved} = 33500\ {}^{kJ}\!/_{mol} \text{ or } 33.5\ MJ/mol$$

b) How many kilojoules of energy must be evolved if you want to get rid of 1.00 kg of this fat by metabolism?
(Value: 2)

37.8 kJ for 1.00 g $\quad \therefore$ 37 800 kJ, or 37.8 MJ , for 1.00 kg of fat

c) An average-sized student burns 365 kJ/h when running at a rate of 10.0 km/h. How many kilometres does the student need to run at this rate to metabolize 1.00 kg of fat?
(Value: 2)

$$\frac{37800\ \text{kJ}}{x} = \frac{365\ \text{kJ}}{1\ \text{hour}} \qquad x = 104\ \text{hours}$$

$$(104\ \text{hours})\left(10.0\ {}^{\text{km}}\!/_{\text{hour}}\right) = 1040\ \text{km}$$

d) Using data from part (a) and that from your tables, calculate the **heat of formation** of fat in kJ/mol.

$C_{57}H_{104}O_6(s) + 80\ O_2(g) \rightarrow 57\ CO_2(g) + 52\ H_2O(l)$

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

ΔH_{rxn} = (57 mol)($CO_2(g)$) + (52 mol)($H_2O(l)$) – (1 mol)($C_{57}H_{104}O_6(s)$) – (80 mol)($O_2(g)$)

ΔHrxn = (57 mol)(-393.5 kJ/mol) + (52 mol)(-285.8 kJ/mol)
– (1 mol)($C_{57}H_{104}O_6(s)$) – (80 mol)(0 kJ/mol)

-33 469.978 kJ = -37 291.1 kJ - (1 mol)($C_{57}H_{104}O_6(s)$)
-33 469.978 kJ + 37 291.1 kJ = -(1 mol)($C_{57}H_{104}O_6(s)$)
+3821.1 kJ = -(1 mol)($C_{57}H_{104}O_6(s)$)
-3821.1 kJ/mol = $C_{57}H_{104}O_6(s)$

ΔHf of $C_{57}H_{104}O_6(s)$ = -3820 kJ/mol

8. 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 $H_2O(l)$, given on page 1 of the Chemistry Data Booklet, to calculate the heat of this reaction measured in kilojoules per mole of HCl(aq)

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

$H = mc\Delta T + mc\Delta T$

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

$H = 1359.8\ J + 1359.8\ J$

$H = 2719.6\ J = 2.7\ kJ$

Mol of HCl = $(C)(V) = (1.00\ mol/L)(0.0500\ L) = 0.05000\ mol$

$$\frac{2.7\ kJ}{x} = \frac{0.0500\ mol}{1.00\ mol}\ x = 54\ kJ$$

$$\Delta H_{rxn} = -54\ kJ/mol\ of\ HCl$$

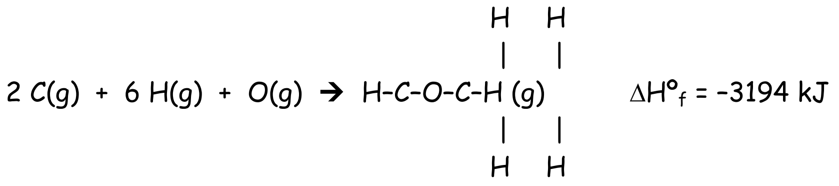
9. We often have to make assumptions when we conduct laboratory experiments. Consider carefully the experiment described in question 10 above, and list four unstated assumptions that must be made in order for one to determine the heat of the reaction.

- We assume that the reaction happens fast enough that no heat is transferred from the system to the environment or vice versa.
- We assume that the densities of HCl(aq), NaOH(aq), and NaCl(aq) all equal 1.00 g/mol.
- 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.
- 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.

10. A knowledge of bond energies is helpful in understanding why some reactions are exothermic while others are endothermic. Notice that bond energies are all positive because energy is always required to break chemical bonds:

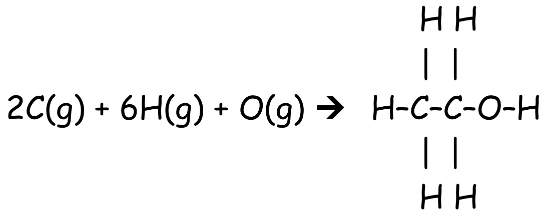
C–H	+413 kJ/mol	C–C	+348 kJ/mol	H–H	+436 kJ/mol
C–O	+358 kJ/mol	O–H	+463 kJ/mol		

These bond energies were used to calculate the enthalpy of formation of $CH_3OCH_3(g)$ from gaseous atoms.



$$2\,C(g) + 6\,H(g) + O(g) \rightarrow \mathrm{H{-}\underset{H}{\overset{H}{C}}{-}O{-}\underset{H}{\overset{H}{C}}{-}H}\,(g) \qquad \Delta H^\circ_f = -3194\text{ kJ}$$

a) Use bond energies to calculate the enthalpy of formation of $CH_3CH_2OH(g)$ from gaseous atoms. Show your calculations.



$$2C(g) + 6H(g) + O(g) \rightarrow \mathrm{H{-}\underset{H}{\overset{H}{C}}{-}\underset{H}{\overset{H}{C}}{-}O{-}H}$$

5 C-H bonds = (5 mol)(-413 kJ/mol) = -2065 kJ
1 C-C bond = (1 mol)(-348 kJ/mol) = -348 kJ
1 C-O bond = (1 mol)(-358 kJ/mol) = -358 kJ
1 O-H bond = (1 mol)(-463 kJ/mol)= $\underline{\underline{-463\text{ kJ}}}$

=-3234 kJ

b) Compare the energy given off in these two equations:

$2\,C(g) + 3\,H2(g) + O(g) \rightarrow CH_3CH_2OH(g)$
$2\,C(g) + 6\,H(g) + O(g) \rightarrow CH_3CH_2OH(g)$

Explain your answer.

Less energy is given off during the formation of $CH_3CH_2OH(g)$ from C(g), $H_2(g)$, and O(g) than is given off during the formation of $CH_3CH_2OH(g)$ from C(g), H(g), and O(g). Energy is required to break the bonds between the hydrogen atoms in H_2 molecules, and this reduced the overall energy given off when $CH_3CH_2OH(g)$ molecules form from C(g), $H_2(g)$, and O(g).

11. Coffee cup calorimeters usually consist of Styrofoam cups with covers, thermometers, and stirrers. This apparatus absorbs a certain amount of heat during calorimetry experiments. Corrections for the heat absorbed by the coffee cup calorimeter is defined as the number of joules required to increase the temperature of the calorimeter by 1°C, measured in J/°C.

    A calorimeter at 20.0°C contains 25.00 g of water also at 20.0°C. When 75.00 g of water at 40.00°C is added to the water in the calorimeter, the highest observed temperature is 34.4°C. Assuming no heat is lost from the calorimeter to the surroundings, what is the heat capacity of this coffee cup calorimeter?

    (Value: 4)

    Heat gained = heat lost

    Cool water + cup = warm water

    $mc\Delta T + c\Delta T = mc\Delta T$

    $(25.00\ g)(4.18\ J/g°C)(14.4°C) + c(14.4°C) = (75.00\ g)(4.18\ J/g°C)(5.6°C)$

    $1540.8\ J + 14.4°C(c) = 1755.6\ J$

    $14.4°C(c) = 250.8\ J$

    c = 17.4166666 = <u>17 J/°C</u>

12. A 2.9 g piece of aluminum, a 6.7 piece of copper, and a 5.0 g piece of iron each absorb the same amount of radiant heat energy from sunlight. Determine which piece of metal will be the warmest. Show your calculations. Use the Chemistry Data Booklet.

    $Q = mc\Delta T \qquad \Delta T = \frac{Q}{mc}$

    Q is the same for all three metals, ∴ let Q = 1

    $$\text{Al} \qquad \Delta T = \frac{Q}{mc} = \frac{1\ J}{(2.9\ g)(0.900\ J/g°C)} = 0.38°C$$

    $$\text{Cu} \qquad \Delta T = \frac{Q}{mc} = \frac{1\ J}{(6.7\ g)(0.385\ J/g°C)} = 0.39°C$$

    $$\text{Fe} \qquad \Delta T = \frac{Q}{mc} = \frac{1\ J}{(5.0\ g)(0.444\ J/g°C)} = 0.45°C$$

    The iron has the largest increase in temperature, ∴ the iron will be the warmest.

13. Coffee cup calorimeters usually consist of Styrofoam cups with covers, thermometers, and stirrers. This apparatus absorbs a certain amount of heat during calorimetry experiments. Corrections for the heat absorbed by the coffee cup calorimeter can be carried out by using the heat capacity of the calorimeter measured in J/°C.

    A 1.00 mol/L solution of NaOH, a 1.00 mol/L solution of HCl, and a coffee cup calorimeter (heat capacity = 109 J/°C) were allowed to stand until the temperature of all three equalled 19.5°C. A 200.0 mL sample of the NaOH was placed in the calorimeter; 200.0 mL of the HCl was added rapidly; and the contents of the calorimeter were mixed. The temperature reached 25.8°C.

    Calculate the heat of neutralization per mole of hydrochloric acid. Assume that the densities of the HCl and NaOH solutions are 1.00 g/mL and that the specific heat capacity of the solution after the reaction is 4.18 J/g°C.

    Heat absorbed = HCl + NaOH + cup
    Heat absorbed= $mc\Delta T + mc\Delta T + c\Delta T$
    Heat = (200.0 g)(4.18 J/g°C)(6.3°C) + (200.0 g)(4.18 J/g°C)(6.3°C)
    + (109 J/°C)(6.3°C)
    Heat absorbed = 5266.8 J + 5266.8 J + 686.7 J
    Heat absorbed = 11 220.3 J = 11.2 kJ

    $\therefore$ heat released by the reaction = -11.2 kJ

    mol HCl = C•V = (1.00 mol/L)(0.2000 L) = 0.200 mol

    $$\frac{0.200\text{ mol}}{1.00\text{ mol}} = \frac{-11.2\text{ kJ}}{x} \qquad x = -56.1\text{ kJ/mol} \qquad \therefore \quad \Delta H_{rxn} = -56.1\text{ kJ/mol}$$

14. A coffee-cup calorimeter normally consists of two nested Styrofoam cups with a lid. A coffee cup calorimeter of this type contains 125.0 g of water at 20.0°C. A 120.0 g sample of copper metal is heated to 98.5°C by putting it into a beaker of hot water. The copper is then put into the calorimeter and the final temperature of the water and copper is 25.8°C.

    a) Calculate the energy change of the copper metal.

    $Q = mc\Delta T$ = (120.0 g)(0.385 J/g°C)(72.7°C) = 3358.74 J
    The copper loses 3360 J.

    b) Calculate the energy change of the water.

    $Q = mc\Delta T$ = (125.0 g)(4.18 J/g°C)(5.8°C) = 3030.5 J
    The water gains 3030 J

c) The difference between the two values is due to the energy lost to the Styrofoam cup. The heat capacity of a calorimeter is the amount of energy required to change the temperature of the apparatus by 1°C. Calculate the heat capacity of the calorimeter in J/°C.

The difference = 3360 J – 3030 J = 330 J

The heat capacity of the calorimeter is: 330 J ÷ 5.8°C = 56.897
= 57 J/°C

d) In a typical coffee-cup calorimeter experiment we assume that the energy change due to the calorimeter is negligible. In this experiment is this a valid assumption? Explain.
(Value: 2)

$$\% \text{ energy lost} = \frac{330\text{ J}}{3360\text{ J}} \times 100 = 9.8\%$$

No, this is not a valid assumption. In this experiment the Styrofoam calorimeter absorbed 330 J, which is 9.8% of the energy released by the copper metal. This would result in too large of a percent error.

15. In a calorimetry experiment the burning of 5.08 g of hexane, $C_6H_{14}(l)$, released enough heat to raise the temperature of 750. g of water from 10.1°C to 82.7°C.

a) Calculate the heat of combustion for hexane expressed in kJ/mol of C_6H_{14}.

$Q = mc\Delta T = (750.\text{ g})(4.18\text{ J/g°C})(72.6°\text{C}) = 227\,601\text{ J} = 228\text{ kJ}$

Moles of hexane = 5.08 g ÷ 86.0 g/mol = 0.05907 = 0.0591 mol

$$\frac{0.0591\text{ mol}}{1\text{ mol}} = \frac{-228\text{ kJ}}{x} \quad x = -3857.868 = -3860\text{ kJ/mol}$$
$$\Delta H_{comb} = -3860\text{ kJ/mol}$$

b) Write a balanced equation for the combustion reaction. Include the ΔH_{comb} in the equation.

$C_6H_{14}(l) + 9\tfrac{1}{2}O_2(g) \rightarrow 6CO_2(g) + 7H_2O(l) \qquad \Delta H_{comb} = -3860\text{ kJ}$

or $2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(l) \qquad \Delta H_{comb} = -7720\text{ kJ}$

16. A chemistry laboratory stocks samples of metals. Unfortunately the labels of some of the "t" metals have fallen off their containers. Some of the samples have been identified; however, the tellurium, terbium, titanium, and tungsten samples have not yet been identified. Use the following data to identify one of the metals.

    When a 60.0 g sample of one of the unlabelled metals at 100.00°C is added to 45.0 g of water at 22.00°C, the final temperature of both the water and the metal is 33.16°C.

    Heat lost by metal = Heat gained by water
    $mc\Delta T = mc\Delta T$

    $(60.0\ g)(c)(66.84°C) = (45.0\ g)(4.18\ J/g°C)(11.16°C)$
    $(4010.4\ g°C)c = 2099.196\ J$
    $c = 2099.196\ J \div 4010.4\ g°C = 0.523\ J/g°C$

    By comparison of the calculated specific heat capacity to those listed in the data booklet, the metal is identified as titanium.

17. The two common sugars, glucose ($C_6H_{12}O_6$), and sucrose ($C_{12}H_{22}O_{11}$), are both carbohydrates.
    a) Using the combustion equations below, calculate the molar enthalpy of combustion for the two sugars. (Value: 4)

    $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$
    $C_{12}H_{22}O_{11}(s) + 12O_2(g) \rightarrow 12CO_2(g) + 11H_2O(l)$

$$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

$\Delta H_{rxn} = \Sigma\Delta H_{f(products)} - \Sigma\Delta H_{f(reactants)}$
$\Delta H_{rxn} = [\ (6\ mol)(CO_2(g)) + (6\ mol)(H_2O(l))\ ]$
$\qquad - [\ (1\ mol)(\ C_6H_{12}O_6(s)) + (6\ mol)(O_2(g))\ ]$
$\Delta H_{rxn} = (6\ mol)(CO_2(g)) + (6\ mol)(H_2O(l)) - (1\ mol)(\ C_6H_{12}O_6(s)) - (6\ mol)(O_2(g))$
$\Delta H_{rxn} = (6\ mol)(-393.5\ kJ/mol) + (6\ mol)(-285.8\ kJ/mol) - (1\ mol)(-1273.1\ kJ/mol)$
$\qquad - (6\ mol)(0\ kJ/mol)$
$\Delta H_{rxn} = -2802.7\ kJ$

$$C_{12}H_{22}O_{11}(s) + 12O_2(g) \rightarrow 12CO_2(g) + 11H_2O(l)$$

$\Delta H_{rxn} = \Sigma\Delta H_{f(products)} - \Sigma\Delta H_{f(reactants)}$
$\Delta H_{rxn} = [\ (12\ mol)(CO_2(g)) + (11\ mol)(H_2O(l))\ ]$
$\qquad - [\ (1\ mol)(C_{12}H_{22}O_{11}(s)) + (12\ mol)(O_2(g))\ ]$
$\Delta H_{rxn} = (12\ mol)(CO_2(g)) + (11\ mol)(H_2O(l)) - (1\ mol)(C_{12}H_{22}O_{11}(s))$
$\qquad - (12\ mol)(O_2(g))$
$\Delta H_{rxn} = (12\ mol)(-393.5\ kJ/mol) + (11\ mol)(-285.8\ kJ/mol)$
$\qquad - (1\ mol)(-2225.5\ kJ/mol) - (12\ mol)(0\ kJ/mol)$
$\mathbf{\Delta H_{rxn} = -5640.3\ kJ}$

b) The average fuel value for carbohydrates is 17 kJ/g.

(i) Calculate the enthalpy of combustion per gram for each sugar and explain how your answers compare to the average fuel value for carbohydrates

$C_6H_{12}O_6$ = 180 g/mol

Enthalpy per gram: $\frac{180\ g}{1\ g} = \frac{-2803\ kJ}{x}$ $x = -15.6\ kJ/g$

$C_{12}H_{22}O_{11}$ = 342 g/mol

Enthalpy per gram: $\frac{342\ g}{1\ g} = \frac{-5640.3\ kJ}{x}$ $x = -16.5\ kJ/g$

Both of these values are close to the average value but are not the same. The more complex carbohydrate has the higher value and there may be a relationship between the size of the molecule and the average fuel value.

(ii) Which of these units, kJ/mol or kJ/g is more likely to be used by chemists? Explain.

A chemist is more likely to use kJ/mol so the value can be used in a chemical equation.

18. Tetrachloroethylene, C_2Cl_4, is used to treat animals infected with hookworms of all types. It can be produced by the reaction:

$$C_2H_2(g) + 3Cl_2(g) \rightarrow C_2Cl_4(g) + 2HCl(g) \qquad \Delta H = -422.1\ kJ$$

a) Using the standard heats of formation, determine the heat of formation of $C_2Cl_4(g)$

$\Delta H_{rxn} = \Sigma\Delta H_{f(products)} - \Sigma\Delta H_{f(reactants)}$

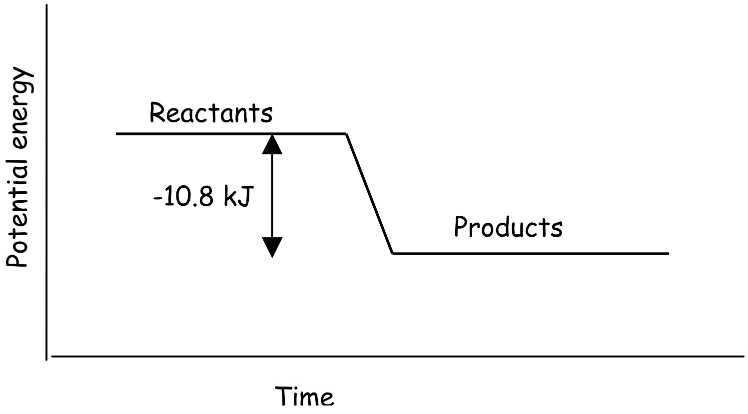
ΔH_{rxn} = [(1 mol)($C_2Cl_4(g)$) + (2 mol)(HCl(g))] – [(1 mol)($C_2H_2(g)$) + (3 mol)($Cl_2(g)$)]

ΔH_{rxn} = (1 mol)($C_2Cl_4(g)$) + (2 mol)(HCl(g)) – (1 mol)($C_2H_2(g)$) - (3 mol)($Cl_2(g)$)

-422.1 kJ = (1 mol)($C_2Cl_4(g)$) + (2 mol)(-92.3 kJ/mol) – (1 mol)(226.7 kJ/mol)
- (3 mol)(o kJ/mol)

ΔHf C_2Cl_4 = -10.8 kJ/mol

b) Sketch an enthalpy diagram for the formation of $C_2Cl_4(g)$ from the elements.(Value: 1)



19. Experienced hikers and campers often make temporary showers using plastic bags. The container is filled from a convenient water source and is left in the sun to absorb heat.

a) If the container holds 30.0 kg of water initially at 9.5°C when collected from a stream, how much energy is absorbed if the final temperature of the water in the "shower" is 35.0°C? (Value: 2)

$Q = mc\Delta T$

$Q = (30.0\ \text{kg})(4.18\ \text{kJ/kg}^\circ\text{C})(25.5^\circ\text{C})$

$Q = 3197.7\ \text{kJ}$

$Q = 3.20 \times 10^3\ \text{kJ}$

b) If a propane stove was used to warm the water instead of sunlight, what mass of propane, $C_3H_8(g)$, must be burned to heat the water from 9.5°C to 35.0°C? (Assume all products of combustion are gases.) (Value: 4)

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$

$\Delta H_{rxn} = \Sigma\Delta H_{f(products)} - \Sigma\Delta H_{f(reactants)}$

$\Delta H_{rxn} = [\ (3\ \text{mol})(CO_2(g)) + (4\ \text{mol})(H_2O(l))\ ] - [\ (1\ \text{mol})(\ C_3H_8(g))$
$+ (5\ \text{mol})(O_2(g))\ ]$

$\Delta H_{rxn} = (3\ \text{mol})(CO_2(g)) + (4\ \text{mol})(H_2O(l)) - (1\ \text{mol})(\ C_3H_8(g)) - (5\ \text{mol})(O_2(g))$

$\Delta H_{rxn} = (3\ \text{mol})(-393.5\ \text{kJ/mol}) + (4\ \text{mol})(-285.8\ \text{kJ/mol}) - (1\ \text{mol})(-103.8\ \text{kJ/mol}) - (6\ \text{mol})(0\ \text{kJ/mol})$

$\Delta H_{rxn} = -2043.9\ \text{kJ}$

$$\frac{2043.9\ \text{kJ}}{3197.7\ \text{kJ}} = \frac{1\ \text{mol}}{x}$$

$$x = 1.56\ \text{mol}\ C_3H_8$$

$$\text{mass} = (\text{mol})(\text{molar mass})$$

$$\text{mass} = (1.56\ \text{mol})(44.09652\ \text{g/mol}) = 69.0\ \text{g}\ \ C_3H_8$$

c) Water is used in several applications where storage of heat energy is a factor. Explain the main thermodynamic advantage of water for this application. (Value: 2)

Water has a specific heat capacity of 4.18 J/g°C. This means it takes a large amount of energy to change the temperature and therefore it has a good heat storage capacity.