Thermochemical equations – heat capacity of water and its uses in determining the ΔH_c Lesson 6

1) Energy output of a reaction can be measured by using this energy to heat a known mass of water and measuring the temperature rise of the water. The heat capacity of water is 4.18 J/g/^{o}C . That is it takes 4.18 Joules of energy to raise one gram of water by one degree Celsius. The formula is used to calculate the energy, in Joules, needed to raise the temperature, Celsius, of a given mass, in grams, of water

$$\text{Energy} = 4.18\ \text{J/g/}^{o}\text{C} \times \text{mass} \times \Delta T$$

Let's try some examples. ***Always give answers to the right number of significant figures***.

a) 100 mL of water was heated from 25.0^{o}C to 28.5 ^{o}C. What amount of energy, in kJ, was used assuming no energy is lost to the environment?

Using the Data booklet we find the density of water at 25^{O}C given at 0.997g/mL.

Step 1 Find the mass of water

=> mass = 0.997 X 100 = 99.7 g

Step 2 Find the energy, in kiloJoules, needed

=> E = 4.18 X 99.7 X (28.5 – 25.0) = 4.18 X 99.7 X 3.5 =1.5 kJ (2 sig figs)

b) 84.7 kJ of heat energy was used to heat 300.0 grams of water at 25.0^{o}C. What was the final temperature of the water assuming no energy was lost to the environment?

Step 1 Find the ΔT

=> ΔT = E / (4.18 X 300.0)

=> ΔT = 84,700 / (4.18 X 300.0) = 67.4 ^{o}C

Step 2 Find the final temperature

=> 67.4+ 25.0 = 92.4^{o}C (3 sig figs)

c) 100.5 kJ of energy is used to heat a body of water from 25.0 ^{o}C to 49.9 ^{o}C. What is the mass of water heated?

Step 1 Find the mass of water

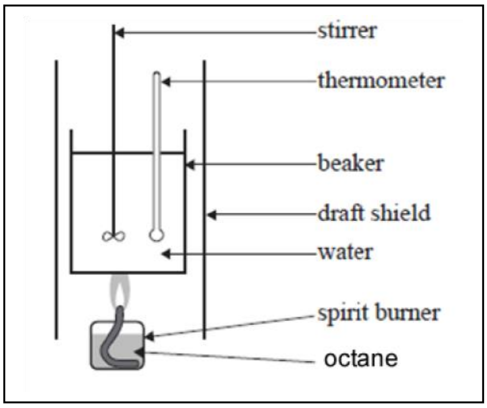
=> mass(g) = E/(4.18 X ΔT) = 100500 /(4.18 X 24.9) = 966g (3 sig figs)

d) What amount of energy, in kJ, was used to raise the temperature of 125 grams of water from 25.1 ^{o}C to 50.0 ^{o}C.

Energy(Joules) = 4.18 J/g/^{o}C X mass X ΔT

=> E = 4.18 X 125 X 24.9 = 13.0 kJ (3 sig figs)

2) Pure octane is placed in a spirit burner and used to heat 100 mL of water. Complete combustion of 0.03510 grams of octane takes place and the temperature of the water rises from 25.0 °C to 28.10 °C.



a) Calculate the amount of energy absorbed by the water.

Energy = 4.18 J/g/°C X mass X ΔT
Step 1 Calculate the mass of water
=> mass(g) = 0.997 X 100 = 99.7g
Step 2 Calculate the energy(J)
=> E = 4.18 X 99.7 X (28.10 – 25.0)
=> E = 4.18 X 99.7 X 3.1 = 1.3 kJ (2 sig figs)

b) Calculate the molar heat of combustion (ΔH_c)

Step 1 calculate the mol of octane
=> 0.03510/114.23 = 0.0003073
Step 2 Molar heat of combustion
=> 1.3kJ/0.0003073 mol = 4.2 X 10^3 *kJ/mol (2 sig figs)*

c) Write a balanced thermochemical equation for the complete combustion of octane.

Octane is a liquid at room at S.L.C. and complete combustion produces CO_2
$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)\ \Delta H = -8.4 \times 10^3$ kJ/mol

d) Compare the ΔH_c of octane derived above in question b) with the Data Sheet. What do you notice? Offer an explanation.

A value of 4,200 kJ/mol is significantly lower than the 5460 kJ/mol given in the Data Booklet. Poor insulation of the apparatus results in significant energy loss.

e) From the data given above and given that the density of octane is 0.702 g/mL calculate the energy density of octane in kJ/L.

Step 1 Calculate the volume of one mol of octane.
=> density = mass(g)/volume(mL)
=> Volume(of one mol) = formula mass/density = 114.23g / 0.702 = 162.7 mL = 0.163L
Step 2 Calculate the energy density
=> 4,200 kJ/mol/0.163L = 2.6 X 10^4 *kJ/L (2 sig figs)*

3) a) 80 mL of water was heated from 25.0°C to 35.5 °C. What amount of energy, in kJ, was used assuming no energy is lost to the environment?

Step 1 – calculate the mass of water
=> 80 X 0.997 = 79.76g
Step 2 – calculate the amount of energy in kJ
=> E = 4.18 X 79.76 X 10.5 = 3.5 kJ or 4 kJ (1 sig fig since 80 mL has only 1 sig fig)

b) 34.2 kJ of heat energy was used to heat 300.0 grams of water at 23.4°C. What was the final temperature of the water assuming no energy was lost to the environment? *50.6°C (3 sig figs)*

c) 43.5 kJ of energy is used to heat a body of water from 25.0 °C to 28.9 °C. What is the mass of water heated?

Step 1 Calculate the ΔT

=> (28.9 - 25.0 = 3.9)

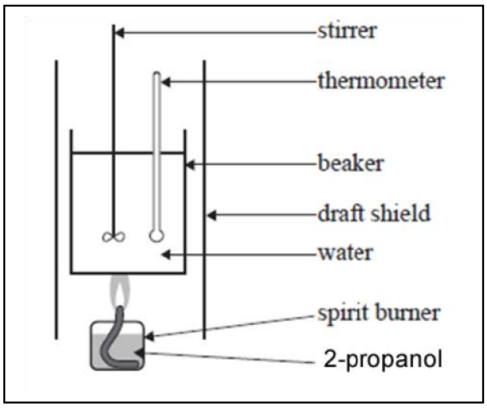
Step 2 calculate the mass in grams

=> 43500 / (4.18 X 3.9) = 2.7 X 10^3 grams (2 sig figs)

d) What amount of energy was used to raise the temperature of 245 grams of water from 25.1 °C to 47.3 °C.

2.27 X 10^4 J (3 sig figs)

4) Pure 2-propanol is placed in a spirit burner and used to heat 100 mL of water. Complete combustion of 0.120 grams of 2-propanol takes place and the temperature of the water rises from 25.0 °C to 31.10 °C.



a) Calculate the amount of energy absorbed by the water.

Energy = 4.18 J/g/°C X mass X ΔT

Step 1 Calculate the mass of water

=> mass(g) = 0.997 X 100 = 99.7g

Step 2 Calculate the energy(J)

=> E = 4.18 X 99.7 X (31.10 – 25.0)

=> E = 4.18 X 99.7 X 6.1 = 2.5 kJ (2 sig figs)

b) Calculate the molar heat of combustion (ΔH_c)

Molar mass of propanol= 60.09 g/mol

Step 1 calculate the mol of propanol

=> 0.120/60.09 = 2.0 X 10^{-3}

Step 2 Molar heat of combustion

=> 2.5/0.0020 = 1.3 X 10^3 kJ/mol (2 sig figs)

c) Write a balanced thermochemical equation for the complete combustion of liquid 2-propanol

$C_3H_8O(l)$ + 4 ½ $O_2(g)$ → $3CO_2(g)$ + $4H_2O(l)$ ΔH = -1.3 X 10^3 kJ/mol

d) The ΔH_c of 2-propanol given in the literature is 2004 kJ/mol. How does this compare with the derived value above in question b) . Explain how this low value can be obtained . *A value of 1,300 kJ/mol is significantly lower than the 2004 kJ/mol given in the literature. Poor insulation of the apparatus results in significant energy loss.*

e) From the data given in the Data Booklet and given that the density of 2-propanol is 0.785 g/mL calculate the true energy density of 2-propanol in kJ/L.

Step 1 Calculate the volume of one mol of 2-propanol.

=> density = mass(g)/volume(mL)

=> Volume(of one mol) = formula mass/density = 60.09g / 0.785 = 76.55 mL = 0.07655L
Step 2 Calculate the energy density
=> 1300 kJ/mol/0.07655L = 1.7 X 10^4 kJ/L (2 sig figs)