Thermochemical equations
Lesson 6B

As of 2024 VCAA has modified the requirement for the expression of ΔH. It is now given in the units of kilojoules. *Eg $3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$ ΔH= -93kJ*

*Refer to the dot point below given in the 2024-2027 VCAA Chemistry study design.*
*comparison of exothermic and endothermic reactions, with reference to bond making and bond breaking, including enthalpy changes (ΔH) measured in kJ, molar enthalpy changes measured in kJ $mol^{-1}$ and enthalpy changes for mixtures measured in kJ $g^{-1}$, and their representations in energy profile diagrams*

1) Below are two thermochemical equations for the phase change of water

   $H_2O(s) \Rightarrow H_2O(l)$ ΔH = +6.00 kJ
   $H_2O(l) \Rightarrow H_2O(g)$ ΔH = +44.01 kJ
   Using the above information explain why the body cools down when we sweat.
   *Heat is removed from the surroundings to change liquid water into gas, thus cooling the skin.*

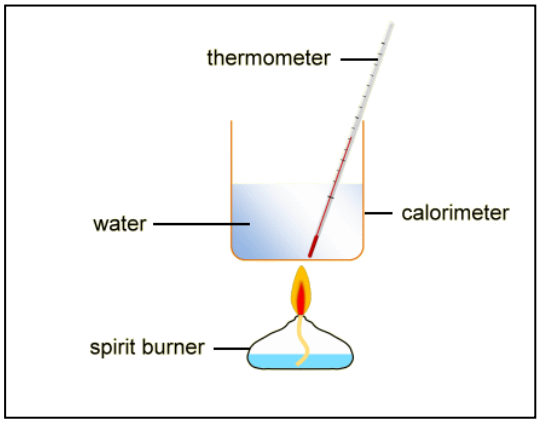
2) Exactly 30.00 g of ethane gas is completely burnt in excess oxygen. During this process it was found that 5,132 kJ of energy was used to break the reactants' bonds and start the reaction process. Upon the formation of products, a total of 6,652 kJ of energy was released.

   Using the experimental data, give the balanced thermochemical equation for the combustion of ethane.

   *$C_2H_6(g) + 3\ ½\ O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$ ΔH = -1520 kJ*

1) Given that the $\Delta H_c$ of glucose is -2816 kJ/mol find the mass of glucose needed to produce 18.40 MJ of energy?
   Step 1 Find the mol of glucose burnt.
   => 18,400 kJ / 2816 kJ = 6.53 mol
   Step 2 Find the mass
   => 6.53 X 180.2 = 1.18 kg

*2)* A spirit burner filled with 2-propanol was used to heat 350.0 mL of water at 12.2°C. The initial mass of the burner was measured at 235.6 g. It was then lit and placed under the beaker of water for 5.00 minutes, the flame was then extinguished and the spirit burner weighed once more.
   a) If the final mass of the burner was 231.6 g and the maximum temperature reached by the water was 94.4°C calculate the $\Delta H_c$ of 2-propanaol. Assume no energy loss from the system.




   *Step 1 Calculate the amount of energy released.*
   *=> E = 4.18 X 350.0 X 82.0 = 1.20 X $10^5$ J*
   *Step 2 Calculate the mol of 2-propanol*
   *=> 4.0/60.0 = 0.0667*
   *Step 3 Calculate the $\Delta H_c$.*
   *=> 120 kJ/ 0.0667 = -1.8 X $10^3$ kJ (2 sig figs)*

b) Write a balanced thermochemical equation for the complete combustion of 2-propanol from the data collected in the experiment above.

*$2C_3H_8O(l) + 9O_2(g) \rightarrow 8H_2O(l) + 6CO_2(g)$*

*A common question is "When do I use liquid water as opposed to water vapour when writing combustion reactions?" Always use liquid state for $H_2O$ if using molar heat of combustion from the data booklet as the molar heat is calculated at SLC. Water is in liquid form at 25 °C unless, of course, the question states that the water is in the gaseous state or indicates in some way that the reaction is not taking place at SLC.*

c) How does the experimentally determined molar heat of combustion for 2-propanol compare with the value given in the literature of -2816 kJ/mol ? Briefly discuss reasons for the discrepancy.

*The experimentally determined $\Delta H_C$ of 2-propanol at $-1.8 \times 10^3$ kJ/mol is significantly lower than the literature value. This is due to the poor insulating properties of the apparatus used. A great deal of heat from the burning of the fuel does not find its way into the water and hence cannot be measured. What is measured through a temperature increase is a fraction of the heat energy generated.*

d) What percentage of the total energy released by the 2-propanol has gone into heating the water?

*(calculated $\Delta H_C$ / actual $\Delta H_C$) X 100 = %*

*The molar heat of combustion is derived from the literature*

=> *$(1.8 \times 10^3 / 2816) \times 100 = 63.2\%$*

e) The density of 2-propanol is 0.785 g/mL. Use this and the experimental value for $\Delta H_C$ (2-propananol), that you obtained in a. above to calculate the energy density of the fuel in $kJ.L^{-1}$.

*Step 1 Find the volume*

*=> V = mass / density = 4.0 / 0.785 g/mL = $5.1 \times 10^{-3}$L*

*Step 2 Find the energy density per litre*

*=> $1.8 \times 10^3 / 5.1 \times 10^{-3}$L = $3.5 \times 10^{-5}$kJ/L (2 sig figs)*