ENTHALPY - A FEW EXAMPLES

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Here are a few more examples of enthalpy calculations.

Example 1. The enthalpy for the reaction in which glucose combines with oxygen to produce carbon dioxide and water is given in the appendix to Schroeder's book as $\Delta H = -1273\text{ kJ mol}^{-1}$. The reaction, which is the source of most energy in mammals, is

$$\mathrm{C_6H_{12}O_6} + 6\mathrm{O_2} \rightarrow 6\mathrm{CO_2} + 6\mathrm{H_2O} \tag{1}$$

The net enthalpy change for this reaction is

$$\Delta H = 6\Delta H_{CO_2} + 6\Delta H_{H_2O} - \Delta H_{\text{glucose}} \tag{2}$$
$$= -6(393.52 + 285.83) + 1273 \tag{3}$$
$$= -2803.1\text{ kJ mol}^{-1} = 670\text{ kcal mol}^{-1} \tag{4}$$

Incidentally, one mole of glucose weighs about $6 \times 12 + 12 + 6 \times 16 = 180$ g so this works out to about 372 kcal per 100 g. The food 'calorie' is actually a kilocalorie so this means that glucose contains about 372 food calories per 100 g.

Example 2. The combustion of one litre of petrol (or 'gasoline' in North America) produces an enthalpy change of -8158 kcal. The current cost (as of July 2015) of petrol here in Scotland is around £1.15 per litre (yes, that's a lot more than in North America), so this works out to

$$\frac{115\text{ pence}}{8158 \times 10^3\text{ cal}} = 1.4 \times 10^{-5}\text{ pence cal}^{-1} \tag{5}$$

To compare this with human (as opposed to car) food, the 'combustion' of corn flakes in the body produces $\Delta H = -100$ kcal per 28 grams. The current cost of a 750 g box of corn flakes here is £1.98 so the cost per calorie is

$$\frac{198}{(750/28) \times 10^5} = 7.4 \times 10^{-5}\text{ pence cal}^{-1} \tag{6}$$

So per calorie, corn flakes are actually about 5 times as expensive as petrol. In North America, where gasoline is a lot cheaper, the difference is probably even greater. Of course, you wouldn't want to drink petrol.

Example 3. The enthalpy of formation of atomic hydrogen gas from molecular hydrogen H_2 is $\Delta H = +217.97$ kJ per mole of *atomic* hydrogen. Since this is positive, energy must be added to molecular hydrogen to dissociate it into atomic hydrogen. To get the enthalpy change per molecule, we double this quantity to get the enthalpy change per mole of molecular hydrogen, then divide by Avogradro's number:

$$\Delta H = \frac{2 \times 217.97 \times 10^3}{6.02 \times 10^{23}} = 7.24 \times 10^{-19} \text{ J} \tag{7}$$

One electron volt is 1.6×10^{-19} J so this is

$$\Delta H = \frac{7.24 \times 10^{-19}}{1.6 \times 10^{-19}} = 4.53 \text{ eV} \tag{8}$$