

Homework #3 Problems (#16-#21)

16. E&R #4.20 [numbers differ from P4.20 2nd ed]

If 3.365 g of ethanol $\text{C}_2\text{H}_5\text{OH}(\ell)$ is burned completely in a bomb calorimeter at 298.15 K, the heat produced is 99.472 kJ.

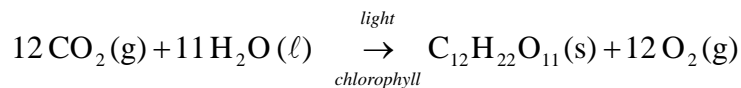
a. Calculate $\Delta H_{\text{combustion}}^0$ for ethanol at 298.15 K.

b. Calculate ΔH_f^0 of ethanol at 298.15 K.

[for part b. use Appendix A (4.1) only; no peeking at A(4.2) !!]

17. (from *Physical Chemistry: Principles and Applications in Biological Sciences*, by Tinoco, Sauer, Wang, Puglisi, published by Prentice Hall).

Photosynthesis can produce 20kg of carbohydrate (e.g. sucrose) per hectare per hour in bright sunlight. Using the following 'model reaction:



a. What is the enthalpy change associated with the production of 1 mol of sucrose from carbon dioxide and water?

b. Assume that photosynthesis can yield 20kg of sucrose per hectare per hour ($5.56 \times 10^{-4} \text{ g m}^{-2} \text{ s}^{-1}$). If sunlight strikes the earth with power 10^3 W m^{-2} , what fraction of this light energy is "stored" as sucrose by photosynthesis (W=watt)?

18. The fermentation of sugar by yeast is one of the oldest chemical processes utilized by *homo sapiens*.



a. Use the data in Appendix A to calculate ΔH^0 for fermentation of 1 mole of α -D-glucose at 25° C

b. If the fermentation reaction was carried out by a thermophilic bacterium at 80° C, what would be ΔH for 1 mole of α -D-glucose (assume temperature independent \overline{C}_p 's)?

19.★ For the combustion of carbon monoxide



The temperature dependent heat capacities of the reactants and products are given by the power series $\bar{C}_p(T) = a + bT + cT^2 + dT^3$ with

(from E&R Appendix A, Table 2.5, 3rd ed):

	a (J mol ⁻¹ K ⁻¹)	b (J mol ⁻¹ K ⁻²)	c (J mol ⁻¹ K ⁻³)	d (J mol ⁻¹ K ⁻⁴)
O₂ (g)	32.83	-0.03633	11.532 × 10 ⁻⁵	-12.194 × 10 ⁻⁸
CO (g)	31.08	-0.01452	3.1415 × 10 ⁻⁵	-1.4973 × 10 ⁻⁸
CO₂ (g)	18.86	0.07937	-6.7834 × 10 ⁻⁵	2.4426 × 10 ⁻⁸

show that (per mole CO):

$$\Delta H(T) = (-2.783 \times 10^5 - 28.635T + 0.0560 T^2 - 5.230 \times 10^{-5} T^3 + 2.509 \times 10^{-8} T^4) \text{ J/mol}$$

20. (ΔH°_f)₂₉₈ for cyclopropane, C₃H₆ (g), is 53.30 kJ/mol.

- Use this value of (ΔH°_f)₂₉₈ to calculate the C—C bond enthalpy for the carbon-carbon bonds in cyclopropane. Assume standard H—H and C—H bond enthalpies of 436 kJ/mole and 413 kJ/mole, respectively, and a $\Delta H_{\text{sublimation}}$ of 717 kJ/mole for C(gr).
- [Tables](#) of bond enthalpies (values averaged over a number of hydrocarbons) give a standard C—C bond enthalpy of 348 kJ/mol. Why does the C—C bond enthalpy [correctly] calculated in part (a) differ from this standard value?
- Would the experimental value of 53.30 kJ/mol be larger or smaller than (ΔH°_f)₂₉₈ calculated using standard bond enthalpies?

21. E&R P4.17d (d only) [same as 2nd ed]