Homework #3
Problems (#16-#21)

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

If 3.365 g of ethanol $C_2H_5OH(\ell)$ is burned completely in a bomb calorimeter at 298.15 K, the heat produced is 99.472 kJ.

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

b. Calculate ΔH^0_f 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:

$$12\,CO_2(g) + 11\,H_2O\,(\ell) \underset{chlorophyll}{\overset{light}{\rightarrow}} C_{12}H_{22}O_{11}(s) + 12\,O_2(g)$$

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 (5.56×10^{-4} g m^{-2} s^{-1}). If sunlight strikes the earth with power 10^3 Wm^{-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*.

$$C_6H_{12}O_6\;(s) \xrightarrow{\text{yeast}} 2C_2H_5OH\,(\ell) + 2CO_2(g)$$

a. Use the data in Appendix A to calculate ΔHº for fermentation of 1 mole of α-D-glucose at 25º C

b. If the fermentation reaction was carried out by a thermophilic bacterium at 80º 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

$CO(g) + \frac{1}{2} O_2(g) \rightarrow CO_2(g)$ $\Delta H^o_{298} = -283$ kJ/mol CO

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})
O_2 (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}
CO_2 (g)	18.86	0.07937	-6.7834×10^{-5}	2.4426×10^{-8}

show that (per mole CO):

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

20. $(\Delta H^o_f)_{298}$ for cyclopropane, C_3H_6 (g) , is 53.30 kJ/mol.

a. Use this value of $(\Delta H^o_f)_{298}$ 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 $\Delta H_{sublimation}$ of 717 kJ/mole for C(gr) .

b. 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?

c. Would the experimental value of 53.30 kJ/mol be larger or smaller than $(\Delta H^o_f)_{298}$ calculated using standard bond enthalpies?

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