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 \( \Delta H^0_{\text{combustion}} \) for ethanol at 298.15 K.
b. Calculate \( \Delta H^0_f \) of ethanol at 298.15 K.

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


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

\[
\text{12 CO}_2(\text{g}) + \text{11 H}_2\text{O}(\ell) \xrightarrow{\text{light chlorophyll}} \text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s}) + \text{12 O}_2(\text{g})
\]

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

\[
\text{C}_6\text{H}_{12}\text{O}_6(\text{s}) \xrightarrow{\text{yeast}} 2\text{C}_2\text{H}_5\text{OH}(\ell) + 2\text{CO}_2(\text{g})
\]

a. Use the data in Appendix A to calculate \( \Delta H^0 \) for fermentation of 1 mole of \( \alpha\)-D-glucose at 25º C
b. If the fermentation reaction was carried out by a thermophilic bacterium at 80º C, what would be \( \Delta H \) for 1 mole of \( \alpha\)-D-glucose (assume temperature independent \( C_p 's \) )?
19. For the combustion of carbon monoxide

\[
\text{CO}(g) + \frac{1}{2} \text{O}_2(g) \rightarrow \text{CO}_2(g) \quad \Delta H^{0}_{298} = -283 \text{ kJ/mol CO}
\]

The temperature dependent heat capacities of the reactants and products are given by the power series \( \overline{C}_p(T) = a + b T + c T^2 + d T^3 \) with

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

<table>
<thead>
<tr>
<th></th>
<th>( a ) (J mol(^{-1}) K(^{-1}))</th>
<th>( b ) (J mol(^{-1}) K(^{-2}))</th>
<th>( c ) (J mol(^{-1}) K(^{-3}))</th>
<th>( d ) (J mol(^{-1}) K(^{-4}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{O}_2 \text{ (g)} )</td>
<td>32.83</td>
<td>-0.03633</td>
<td>11.532 \times 10^{-5}</td>
<td>-12.194 \times 10^{-8}</td>
</tr>
<tr>
<td>( \text{CO} \text{ (g)} )</td>
<td>31.08</td>
<td>-0.01452</td>
<td>3.1415 \times 10^{-5}</td>
<td>-1.4973 \times 10^{-8}</td>
</tr>
<tr>
<td>( \text{CO}_2 \text{ (g)} )</td>
<td>18.86</td>
<td>0.07937</td>
<td>-6.7834 \times 10^{-5}</td>
<td>2.4426 \times 10^{-8}</td>
</tr>
</tbody>
</table>

dshow 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. \( (\Delta H^0_f)_{298} \) for cyclopropane, \( \text{C}_3\text{H}_6 \text{ (g)} \), is 53.30 kJ/mol.

a. Use this value of \( (\Delta H^0_f)_{298} \) 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).

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 \( (\Delta H^0_f)_{298} \) calculated using standard bond enthalpies?

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