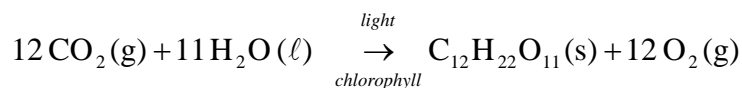


### Homework #4 Problems (#17-#21)

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:



- What is the enthalpy change associated with the production of 1 mol of sucrose from carbon dioxide and water?
  - 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 ( $W=\text{watt}$ )?
18. The fermentation of sugar by yeast is one of the oldest chemical processes utilized by *homo sapiens*.



- Use the data in Appendix A to calculate  $\Delta H^\circ$  for fermentation of 1 mole of  $\alpha$ -D-glucose at  $25^\circ \text{C}$
- If the fermentation reaction was carried out by a thermophilic bacterium at  $80^\circ \text{C}$ , what would be  $\Delta H$  for 1 mole of  $\alpha$ -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  $\overline{C}_p(T) = a + bT + cT^2 + dT^3$  with

(from E&R Appendix A, Table 2.5, 3<sup>rd</sup> ed):

	a (J mol <sup>-1</sup> K <sup>-1</sup> )	b (J mol <sup>-1</sup> K <sup>-2</sup> )	c (J mol <sup>-1</sup> K <sup>-3</sup> )	d (J mol <sup>-1</sup> K <sup>-4</sup> )
O <sub>2</sub> (g)	32.83	-0.03633	$11.532 \times 10^{-5}$	$-12.194 \times 10^{-8}$
CO (g)	31.08	-0.01452	$3.1415 \times 10^{-5}$	$-1.4973 \times 10^{-8}$
CO <sub>2</sub> (g)	18.86	0.07937	$-6.7834 \times 10^{-5}$	$2.4426 \times 10^{-8}$

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.  $(\Delta H^{\circ}_f)_{298}$  for cyclopropane,  $C_3H_6(g)$ , is 53.30 kJ/mol.
- Use this value of  $(\Delta H^{\circ}_f)_{298}$  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  $(\Delta H^{\circ}_f)_{298}$  calculated using standard bond enthalpies?
21. E&R P4.17d (d only)