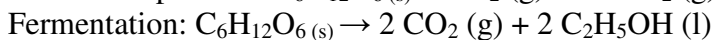
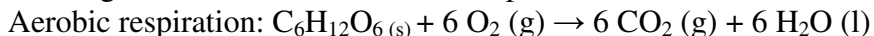


## Workshop key for May 6 Chemistry workshop

1. Some counterexamples include evaporation, sublimation, or melting, all of which are endothermic. The chemical cold pack, in which the dissolution of ammonium nitrate is endothermic is another example.
2. A. Boiling, the conversion of liquid to gas, results in an entropy increase for the substance in question. B. We can consider the boiling to be a reversible process and use the definition  $\Delta S = q_{\text{reversible}}/T$ . Here,  $q$  is the enthalpy of vaporization (71.8 kJ/mole) and  $T$  is the absolute temperature ( $64.7 + 273.1 = 337.8$  K).  $\Delta S = 71,800/337.8 = 213$  J/mole-K or 0.213 kJ/mole-K C. To find the heat needed to vaporize a given mass of methanol we take the heat of vaporization per mole times the number of moles. For methanol 202 g is  $202/32 = 6.31$  moles. The heat needed is thus  $6.31$  moles \* 71.8 kJ/mole = 453 kJ
3. In each case here the standard entropy change is the standard entropy of the products minus the standard entropy of the reactants. Make sure to use the correct state of matter, as solid, gas, liquid or solution will have different thermodynamic value. For this problem all units are J/mole-K.
  - A.  $2 * S^0(\text{NH}_3, \text{g}) - [S^0(\text{N}_2\text{H}_4, \text{g}) + S^0(\text{H}_2, \text{g})] = 2*(192.5) - [238.5 + 130.58] = 15.92$  (slightly positive, and the sign somewhat hard to predict. It is a small entropy change, which fits a reaction in which 2 reactant gas molecules make 2 product gas molecules.)
  - B.  $2 * S^0(\text{AlCl}_3, \text{s}) - [2 * S^0(\text{Al}, \text{s}) + 3 * S^0(\text{Cl}_2, \text{g})] = 2*109.3 - [2*28.32 + 3*222.96] = -506.92$  J/mole-K. (A large negative entropy change, which fits as several gas molecules are converted into a few solid molecules.)
  - C.  $[S^0(\text{MgCl}_2, \text{s}) + 2 * S^0(\text{H}_2\text{O}, \text{l})] - [2 * S^0(\text{Mg}(\text{OH})_2, \text{s}) + 2 * S^0(\text{HCl}, \text{g})] = [89.6 + 2 * 69.91] - [63.24 + 2 * 186.69] = 229.42 - 436.62 = -207$  J/mole-K. (Again, converting some gas into liquid and solid, which would reduce entropy of the system,)
4. A  $\Delta H^0 = -305.3$  kJ/mole ;  $\Delta G^0 = -259.0$  kJ/mole,  $\Delta S^0 = -155.21$  J/mole-K Check:  $\Delta H - T\Delta S$  at 298 =  $-305.3 - 298(-0.15521)$  kJ/mole-K =  $-259.0$  kJ/mol: the values are consistent. The most common problem here is not having the entropy, enthalpy, and free energy in the same energy units. Note that some values are in J and some are in kJ.
  - B.  $\Delta H^0 = 178.1$  kJ/mole,  $\Delta G^0 = 130.19$  kJ/mole  $\Delta S^0 = 160.47$  J/mole -K.  $\Delta H - T\Delta S$  at 298 =  $178.1 - 298(0.16047)$  kJ/mole-K =  $130.28$  kJ/mole. Close fit, probably round off errors in calculation.)
  - C.  $\Delta H^0 = -1453.12$  kJ/mole,  $\Delta G^0 = -1404.86$  kJ/mole  $\Delta S^0 = -161.8$  J/mole -K.  $\Delta H - T\Delta S = 1453.12 - 298(-0.1618) = -1404.9$  kJ/rxn. Fits well.
5. A.  $\Delta H^0 = +18.7$  kJ/mole  $\Delta S^0 = 240$  J/mol-K. At standard conditions  $\Delta G = 18.7 - 298(0.240) = -52$  kJ/mole  $\rightarrow$  thus spontaneous.
  - B. Find the  $T$  at which  $\Delta G$  is equal to 0. Below this temperature not spontaneous.  $\Delta G = 0$  at 78 K, so any temperature above this is spontaneous.
  - C. We had to assume that the values of  $S$  and  $H$  from the table were not significantly different at 78 K than they were at standard conditions (298 K). In this problem this is a large  $T$  difference, so it may not be such a good assumption.
6. Although the entropy of the system decreases, this reaction also has a large negative enthalpy change-a lot of heat is released to the surroundings. This gives the

surroundings a large positive entropy change, larger in magnitude than the negative entropy change of the system, so the sum  $\Delta S_{\text{system}} + \Delta S_{\text{surroundings}} > 0$ .

7. Using the tabulated data for the two processes:



- A. Maximum amount of work is the change in free energy ( $\Delta G$ ) for each reaction, for aerobic respiration this is (at standard conditions)  $-2878 \text{ kJ/mole glucose}$ , so the system could do  $2878 \text{ kJ}$  of work on the surroundings per mole glucose. For fermentation the free energy change is  $-227.7 \text{ kJ/mole glucose}$  (Note that this indicates that the maximum energy yield from fermentation is about 8 % that of aerobic respiration.)
- B. For this calculate the enthalpy change for fermentation.  $\Delta H^0$  for fermentation is ( $-85.6 \text{ kJ/mole glucose}$ ). Since this system value is negative, the reaction is exothermic and this amount of heat must be removed per mole of glucose fermented.