

Entropy and Gibbs Free Energy

Entropy

1. Calculate the change in entropy of a large vat of molten copper when 50 J of energy is removed reversible from it as heat at 1100 °C.

$$\Delta S = \frac{q_{\text{rev}}}{T} = \frac{-50 \text{ J}}{(1100 + 273) \text{ K}} = -0.0364 \frac{\text{J}}{\text{K}}$$

b/c the 50J is removed (heat out of system)

2. Calculate the change in entropy of 1.0 L of water at 0 °C when it absorbs 235 J of energy from a heater. (b) If the 1.0 L of water is at 99 °C, what is its entropy change? (c) Explain any difference in entropy change.

1L H₂O liquid = 1000 mL = 1000 g

$$(a) \Delta S = \frac{q_{\text{rev}}}{T} = \frac{235 \text{ J}}{(0 + 273) \text{ K}} = 0.861 \frac{\text{J}}{\text{K}}$$

$$(b) \Delta S = \frac{235 \text{ J}}{(99 + 273) \text{ K}} = 0.632 \frac{\text{J}}{\text{K}}$$

- (c) The entropy change is smaller at higher temperatures b/c the matter is already more chaotic.
3. What is the standard entropy of vaporization of acetone at its normal boiling point of 56.7 °C? The standard enthalpy of vaporization for acetone is 29.1 kJ/mol.

$$\Delta S_{\text{vap}} = \frac{\Delta H_{\text{vap}}}{T_{\text{BP}}} = \frac{29.1 \text{ kJ/mol}}{(56.7 + 273) \text{ K}} = 0.0883 \frac{\text{kJ}}{\text{mol} \cdot \text{K}} = 88.3 \frac{\text{J}}{\text{mol} \cdot \text{K}}$$

4. Calculate the entropy change for the vaporization of 50.0 g of ethanol at 351.5 K. $\Delta H_{\text{vap}}^{\circ}(\text{ethanol}) = 43.5 \text{ kJ/mol}$.

$$n(\text{ethanol}) = \frac{m}{\text{MW}} = \frac{50 \text{ g}}{(2 \times 12 + 6 \times 1 + 1 \times 16) \text{ g/mol}} = 1.087 \text{ mol}$$

$$\Delta S = \frac{\Delta H_{\text{vap}}^{\circ} \times n}{T} = \frac{43.5 \frac{\text{kJ}}{\text{mol}} \times 1.087 \text{ mol}}{351.5 \text{ K}} = 0.1345 \frac{\text{kJ}}{\text{K}} = 134.5 \frac{\text{J}}{\text{K}}$$

(remember these values are per 50 g of ethanol)

5. Calculate the entropy of fusion of mercury at its freezing point (234.3 K). $\Delta H_{\text{fus}}^{\circ}(\text{mercury}) = 2.292 \text{ kJ/mol}$.

$$\Delta S_{\text{fus}} = \frac{\Delta H_{\text{fus}}}{T_{\text{mp}}} = \frac{2.292 \text{ kJ/mol}}{234.3 \text{ K}} = 9.78 \frac{\text{J}}{\text{mol} \cdot \text{K}}$$

6. The entropy of vaporization of benzene is approximately $85 \text{ J}/(\text{K}\cdot\text{mol})$ and the enthalpy of vaporization of benzene is $30 \text{ kJ}/\text{mol}$. What is the boiling point of benzene?

$$\Delta S_{\text{vap}} = \frac{\Delta H_{\text{vap}}}{T_{\text{BP}}} \Rightarrow T_{\text{BP}} = \frac{\Delta H_{\text{vap}}}{\Delta S_{\text{vap}}} = \frac{30 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}}}{85 \frac{\text{J}}{\text{K}\cdot\text{mol}}}$$

$$\Rightarrow \boxed{T_{\text{BP}} = 352.9 \text{ K} = 79.9^\circ\text{C}}$$

7. Suppose that 104 g of cyclohexane at 17°C is mixed with 234 g of cyclohexane at 77°C at constant atmospheric pressure in a thermally insulated vessel. What is the ΔS_{sys} for the process? The specific heat capacity for cyclohexane is, $C_{p,m} = 156.5 \text{ J}/(\text{K}\cdot\text{mol})$

$$77^\circ\text{C} = 350 \text{ K} \quad + \quad 17^\circ\text{C} = 290 \text{ K} \quad m(\text{1 mol cyclohexane}) = 6 \times 12 + 12 \times 1 = 84 \text{ g/mol}$$

[1] We first need to find T_f which will be the same for the "hot" and "cool" side by setting $q_{\text{hot}} = -q_{\text{cool}}$. We can do this b/c it is done in a thermally isolated vessel so all the heat released by the "hot" side is absorbed by the "cool" side.

[2] We need to change the molar heat capacity given to the (mass) heat capacity:

$$C_p = 156.5 \frac{\text{J}}{\text{K}\cdot\text{mol}} \times \frac{1 \text{ mol}}{84 \text{ g}}$$

$$= 1.863 \frac{\text{J}}{\text{g}\cdot\text{K}}$$

$$q_h = -q_c \Rightarrow m_h C_{p(\text{cyclohexane})} (T_f - T_{i,h}) = -m_c C_{p(\text{cyclohexane})} (T_f - T_{i,c})$$

$$\Rightarrow T_f m_h + T_f m_c = T_{i,c} m_c + T_{i,h} m_h$$

$$\Rightarrow T_f = \frac{T_{i,c} m_c + T_{i,h} m_h}{m_h + m_c} = \frac{(290 \text{ K})(104 \text{ g}) + (350 \text{ K})(234 \text{ g})}{(104 \text{ g}) + (234 \text{ g})}$$

[3] With T_f we can now calculate ΔS of the "hot" & "cold" side using the eq.:

$$\Delta S_c = n_c C_p \ln\left(\frac{T_f}{T_{i,c}}\right) = (104 \text{ g}) \left(1.863 \frac{\text{J}}{\text{g}\cdot\text{K}}\right) \ln\left(\frac{331.5 \text{ K}}{290 \text{ K}}\right) = 25.91 \frac{\text{J}}{\text{K}}$$

$$\Delta S_h = (234 \text{ g}) \left(1.863 \frac{\text{J}}{\text{g}\cdot\text{K}}\right) \ln\left(\frac{331.5 \text{ K}}{350 \text{ K}}\right) = -23.67 \frac{\text{J}}{\text{K}}$$

[4] We add them together to get $\Delta S_{\text{total system}} = \Delta S_c + \Delta S_h = 25.91 - 23.67 = 2.24 \frac{\text{J}}{\text{K}}$

8. Using data of S_m° from the Appendix in your book, calculate the standard entropy change for each of the following reactions at 25°C . For each reaction, interpret the sign and magnitude of the reaction entropy.



$$\Delta S_r^\circ = \sum n S_m^\circ(\text{products}) - \sum n S_m^\circ(\text{reactants})$$

$$\Delta S_r^\circ = S_m^\circ(\text{CO}_2) + S_m^\circ(\text{CaO}) - S_m^\circ(\text{CaCO}_3)$$

$$\boxed{\Delta S_r^\circ = (213.74 \frac{\text{J}}{\text{K}\cdot\text{mol}}) + (39.75 \frac{\text{J}}{\text{K}\cdot\text{mol}}) - (92.9 \frac{\text{J}}{\text{K}\cdot\text{mol}}) = 160.59 \frac{\text{J}}{\text{K}\cdot\text{mol}}}$$



$$\Delta S_r^\circ = 3 S_m^\circ(\text{KClO}_4) + S_m^\circ(\text{KCl}) - 4 S_m^\circ(\text{KClO}_3)$$

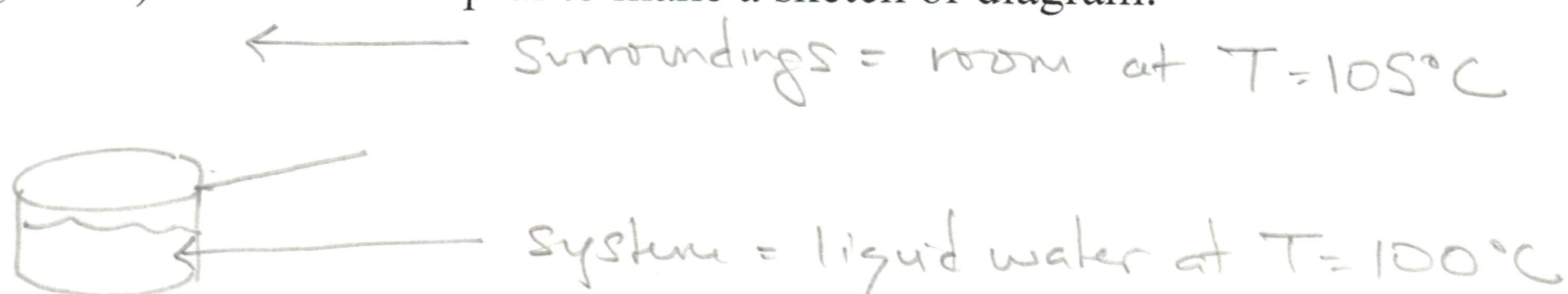
$$\boxed{\Delta S_r^\circ = 3(151.0 \frac{\text{J}}{\text{K}\cdot\text{mol}}) + (82.59 \frac{\text{J}}{\text{K}\cdot\text{mol}}) - 4(143.1 \frac{\text{J}}{\text{K}\cdot\text{mol}}) = -36.81 \frac{\text{J}}{\text{K}\cdot\text{mol}}}$$

The Familiar Process of Water Boiling

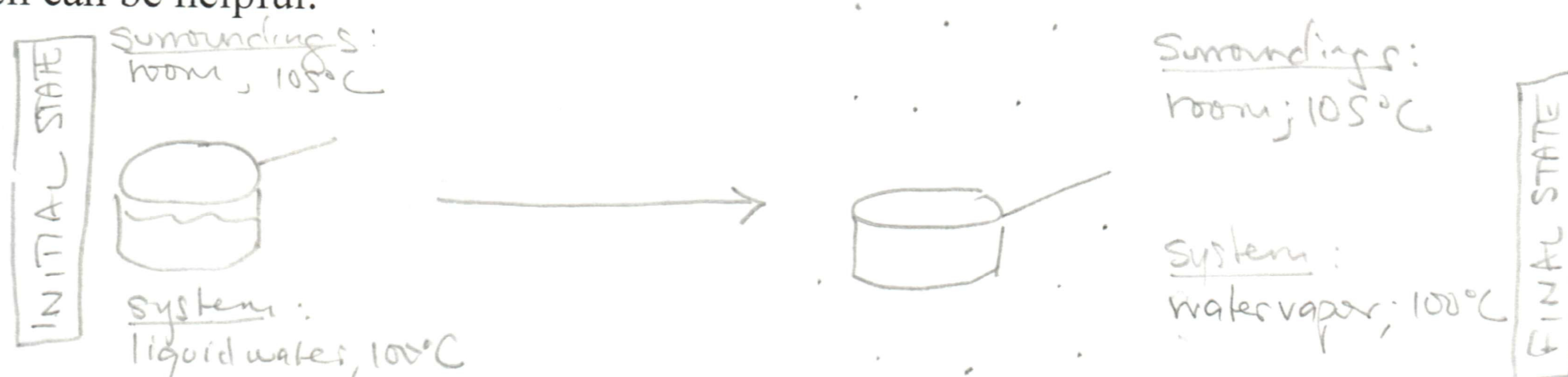
Imagine you start with a pot of 54 g of water. The initial temperature of this mixture is 100 °C. If you leave this pot in a room that has a constant temperature of 105 °C, the water will *spontaneously* boil and you will end up with all water vapor at 100 °C.

Here we will calculate the entropy change for the process of the water boiling (note: after the water boils the temperature of the resulting water vapor will rise until it is the same temperature as the room. We are just looking at the first step where the liquid water boils).

1. First identify what the system and surroundings are. (what are they composed of, what is their temperature, etc...) Often it is helpful to make a sketch or diagram.



2. Now identify what the initial and final states of the system are. Again a diagram, equation, or sketch can be helpful.



3. Now look at the energy change for this process. Does energy flow into or out of the system? If so, as heat or work?

• energy flow into the system in the form of heat

* here we ignore the energy flow out of the system in the form of work from the gas formation/expansion.

4. Given that $\Delta_{\text{vap}}H^\circ = 40.7 \text{ kJ mol}^{-1}$, what is the change in enthalpy for system for this process (in Joules)?

$$\boxed{1} \quad n(\text{H}_2\text{O}) = \frac{54\text{g}}{[(1 \times 16) + (2 \times 1)] \frac{\text{g}}{\text{mol}}} = 3 \text{ moles}$$

$$\boxed{2} \quad \Delta H = \Delta H_{\text{vap}} \times n_{\text{H}_2\text{O}} = 40.7 \frac{\text{kJ}}{\text{mol}} \times 3 \text{ mol}$$

$$\boxed{\Delta H = 122.1 \text{ kJ}}$$

5. Given that this process is at constant temperature, $q = \Delta H$. This is the *reversible heat*. What then is the change in entropy for the system?

$$\boxed{\Delta S = \frac{q_{\text{rev}}}{T} = \frac{\Delta H}{T} = \frac{122.1 \text{ kJ}}{(100 + 273) \text{ K}} = 0.3273 \frac{\text{kJ}}{\text{K}} = 327.3 \frac{\text{J}}{\text{K}}}$$

6. What is the heat from the perspective of the surroundings? (did the energy go up? Down? How much?)

$$q_{\text{surr}} = -q_{\text{sys}} = -\Delta H_{\text{vap}} \times n = -40.7 \frac{\text{kJ}}{\text{mol}} \times 3 \text{ mol} = -122.1 \text{ kJ}$$

7. What is the entropy change for the surroundings?

$$\Delta S_{\text{surr}} = \frac{q_{\text{surr}}}{T} = \frac{-122.1 \text{ kJ}}{373 \text{ K}} = -0.3230 \frac{\text{kJ}}{\text{K}} = -323.0 \frac{\text{J}}{\text{K}}$$

8. What is the entropy change for the universe for this process? Does this make sense with what physically is happening?

$$\Delta S_{\text{universe}} = \Delta S_{\text{surr}} + \Delta S_{\text{sys}} = 327.3 \frac{\text{J}}{\text{K}} - 323.0 \frac{\text{J}}{\text{K}} = +4.3 \frac{\text{J}}{\text{K}}$$

$\Delta S_{\text{universe}}$ is positive which corresponds to a spontaneous process. This makes sense b/c the boiling of water is spontaneous in these conditions.

9. The answer to part 8 is extensive (this for the 54 g of liquid water that boils in this process). What is the entropy change per mole of liquid water that boils under these conditions?

The ΔS_{uni} calculated in 8 was for 3 moles of H_2O . So we need to divide our answer in 8 by 3 to get the $\Delta S/\text{mole}$ of liquid water under these conditions.

$$\Delta S_{\text{uni},m} = +4.3 \frac{\text{J}}{\text{K}} \div 3 = +1.43 \frac{\text{J}}{\text{K}}$$

10. Redo steps 7 and 8 but with a room that has a constant temperature of 95 °C. Explain any observed differences or similarities.

(#7) $\Delta S_{\text{surr}} = \frac{-122.1 \text{ kJ}}{368 \text{ K}} = -0.3318 \frac{\text{kJ}}{\text{K}} = -331.8 \frac{\text{J}}{\text{K}}$

(#8) $\Delta S_{\text{uni}} = 327.3 \frac{\text{J}}{\text{K}} - 331.8 \frac{\text{J}}{\text{K}} = -4.5 \frac{\text{J}}{\text{K}}$

Under these conditions, ΔS_{uni} is negative which corresponds to a non-spontaneous change which is what we would observe: Liquid water does NOT boil spontaneously when placed in a room of temperature 95 °C b/c it is below water's boiling pt.

Gibb's Free Energy

$$\text{Gibb's Free Energy Equation: } \Delta G = \Delta H - T\Delta S$$

1. Under what conditions is the equation above true?

- constant temperature
- constant pressure

2. If ΔH of a process is positive and ΔS of that same process is positive, what can you conclude about the spontaneity of the process?

a. The process is spontaneous. How do you know that?

b. The process is not spontaneous. How do you know that?

c. There is not enough information to conclude. If this is the case, what information would you need to be able to make a conclusion?

the temperature

3. If ΔH of a process is negative and ΔS of that same process is negative, what can you conclude about the spontaneity of the process?

a. The process is spontaneous. How do you know that?

b. The process is not spontaneous. How do you know that?

c. There is not enough information to conclude. If this is the case, what information would you need to be able to make a conclusion?

the temperature

4. If ΔH of a process is positive and ΔS of that same process is negative, what can you conclude about the spontaneity of the process?

a. The process is spontaneous. How do you know that?

b. The process is not spontaneous. How do you know that?

b/c ΔG is positive

c. There is not enough information to conclude. If this is the case, what information would you need to be able to make a conclusion?

5. If ΔH of a process is negative and ΔS of that same process is positive, what can you conclude about the spontaneity of the process?

a. The process is spontaneous. How do you know that?

b/c ΔG is negative

b. The process is not spontaneous. How do you know that?

c. There is not enough information to conclude. If this is the case, what information would you need to be able to make a conclusion?

6. Calculate the standard enthalpy change, entropy change, and free-energy change at 298°C for each of the following reactions by using data in the Appendix in the back of your book. For each case, confirm that the value obtained from the free energies of formation is the same as that obtained by using the relation $\Delta G_r^\circ = \Delta H_r^\circ - T \Delta S_r^\circ$.

The oxidation of magnetite to hematite: $2\text{Fe}_3\text{O}_4(\text{s}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow 3\text{Fe}_2\text{O}_3(\text{s})$

$$\Delta H_r^\circ = \sum n \Delta H_f^\circ(\text{products}) - \sum n \Delta H_f^\circ(\text{reactants})$$

$$= 3 \Delta H_f^\circ(\text{Fe}_2\text{O}_3(\text{s})) - [2 \Delta H_f^\circ(\text{Fe}_3\text{O}_4(\text{s})) + \frac{1}{2} \Delta H_f^\circ(\text{O}_2(\text{g}))]$$

$$= 3(-824.2) - [2(-1118.4) + \frac{1}{2}(0)] = -235.8 \text{ kJ/mol}$$

$$\Delta S_r^\circ = \sum n S^\circ(\text{products}) - \sum n S^\circ(\text{reactants})$$

$$= 3 S^\circ(\text{Fe}_2\text{O}_3(\text{s})) - [2 S^\circ(\text{Fe}_3\text{O}_4(\text{s})) + \frac{1}{2} S^\circ(\text{O}_2(\text{g}))]$$

$$= 3(87.4) - [2(146.4) + \frac{1}{2}(205.14)] = -133.3 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta G_r^\circ = \Delta H_r^\circ - T \Delta S_r^\circ = -235.8 \text{ kJ/mol} - (298 \text{ K})(-133.3 \frac{\text{kJ}}{\text{mol}}) = -196.1 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta G_r^\circ = \sum n \Delta G_f^\circ(\text{products}) - \sum n \Delta G_f^\circ(\text{reactants})$$

$$= 3 \Delta G_f^\circ(\text{Fe}_2\text{O}_3(\text{s})) - [2 \Delta G_f^\circ(\text{Fe}_3\text{O}_4(\text{s})) + \frac{1}{2} \Delta G_f^\circ(\text{O}_2(\text{g}))]$$

$$= 3(-742.2) - [2(-1015.4) + \frac{1}{2}(0)] = -195.8$$

The dimerization of NO_2 : $2\text{NO}_2(\text{g}) \rightarrow \text{N}_2\text{O}_4(\text{g})$

$$\Delta H_r^\circ = \Delta H_f^\circ(\text{N}_2\text{O}_4(\text{g})) - 2 \Delta H_f^\circ(\text{NO}_2(\text{g})) = +9.16 - 2(+33.18) = -57.2 \frac{\text{kJ}}{\text{mol}}$$

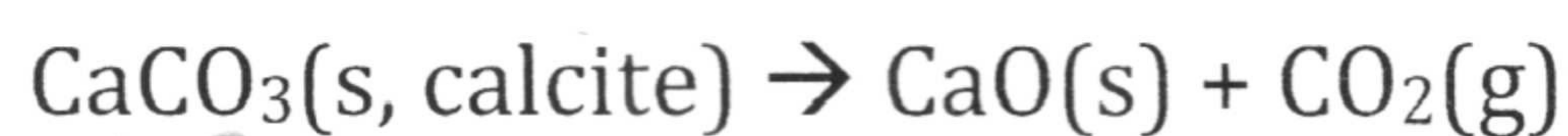
$$\Delta S_r^\circ = S^\circ(\text{N}_2\text{O}_4(\text{g})) - 2 S^\circ(\text{NO}_2(\text{g})) = 304.29 - 2(240.06) = -75.83 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta G_r^\circ = \Delta H_r^\circ - T \Delta S_r^\circ = -57.2 \frac{\text{kJ}}{\text{mol}} - (298 \text{ K})(-75.83 \frac{\text{kJ}}{\text{mol}}) = -4.8 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta G_r^\circ = \Delta G_f^\circ(\text{N}_2\text{O}_4(\text{g})) - 2 \Delta G_f^\circ(\text{NO}_2(\text{g})) = +97.89 - 2(+51.31) = -4.73$$

Both methods of calculating ΔG_r° gives very similar values as expected.

7. Use the standard free energies of formation in the Appendix in the back of your book to calculate ΔG_r° for each of the following reactions at 25 °C. Comment on the spontaneity of each reaction under the standard conditions at 25 °C.

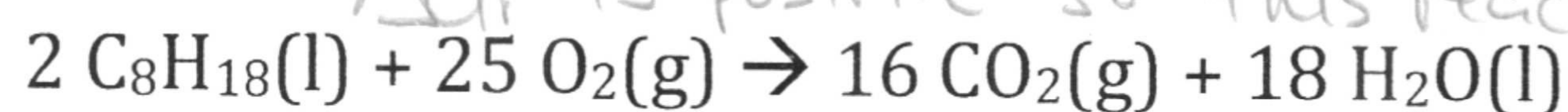


$$\Delta G_r^\circ = \sum n \Delta G_f^\circ(\text{products}) - \sum n \Delta G_f^\circ(\text{reactants})$$

$$\Delta G_r^\circ = \Delta G_f^\circ(\text{CaO}(\text{s})) + \Delta G_f^\circ(\text{CO}_2(\text{g})) - \Delta G_f^\circ(\text{CaCO}_3(\text{s, calcite}))$$

$$\Delta G_r^\circ = -604.03 + (-394.36) - (-1128.8) = +130.41 \text{ kJ/mol}$$

ΔG_r° is positive so this reaction is NOT spontaneous.



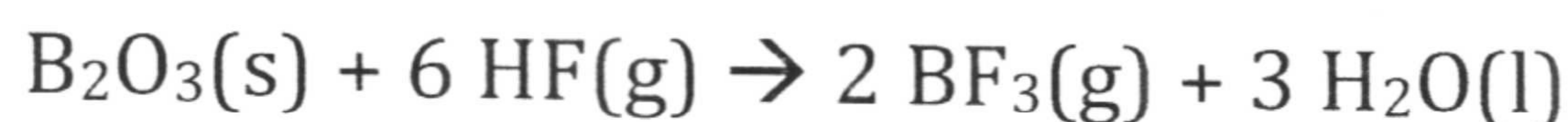
$$\Delta G_r^\circ = 16 \Delta G_f^\circ(\text{CO}_2(\text{g})) + 18 \Delta G_f^\circ(\text{H}_2\text{O}(\text{l})) - [2 \Delta G_f^\circ(\text{C}_8\text{H}_{18}(\text{l})) + 25 \Delta G_f^\circ(\text{O}_2(\text{g}))]$$

$$\Delta G_r^\circ = 16 \times (-394.36) + 18 \times (-237.13) - [2(+61.4) + 25(0)] = -10590.9 \text{ kJ/mol}$$

ΔG_r° is negative so this is a spontaneous reaction

8. Assume that ΔH_r° and ΔS_r° are independent of temperature and use data in the Appendix in the back of your book to calculate ΔG_r° for each of the following reactions at 80°C . Over what temperature range will each reaction be spontaneous under standard conditions?

$$80^\circ\text{C} + 273 = 353 \text{ K}$$



$$\Delta H_r^\circ = 2 \Delta H_f^\circ(\text{BF}_3(\text{g})) + 3 \Delta H_f^\circ(\text{H}_2\text{O}(\text{l})) - [\Delta H_f^\circ(\text{B}_2\text{O}_3(\text{s})) + 6 \Delta H_f^\circ(\text{HF}(\text{g}))]$$

$$= 2(-1137.0) + 3(-285.83) - [(-1272.8) + 6(-271.1)] = -232.09 \text{ kJ/mol}$$

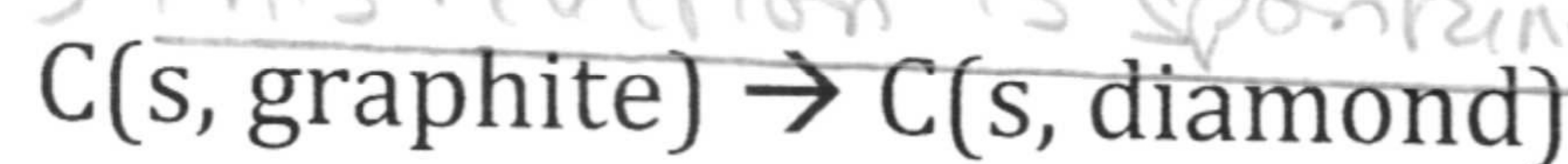
$$\Delta S_r^\circ = 2 S_m^\circ(\text{BF}_3(\text{g})) + 3 S_m^\circ(\text{H}_2\text{O}(\text{l})) - [S_m^\circ(\text{B}_2\text{O}_3(\text{s})) + 6 S_m^\circ(\text{HF}(\text{g}))]$$

$$= 2(254.12) + 3(69.91) - [(53.97) + 6(173.78)] = -378.68 \text{ J/mol}$$

$$\Delta G_r^\circ = -232.09 \frac{\text{kJ}}{\text{mol}} - 353 \text{ K} \left(-0.37868 \frac{\text{kJ}}{\text{mol}} \right) = -98.4 \frac{\text{kJ}}{\text{mol}}$$

The reaction is spontaneous if $T < \frac{\Delta H_r^\circ}{\Delta S_r^\circ} = \frac{-232.09 \frac{\text{kJ}}{\text{mol}}}{-0.37868 \frac{\text{kJ}}{\text{mol}}} = 612.9 \text{ K} = 339.9^\circ\text{C}$

So, this reaction is spontaneous below 612.9 K .



$$\Delta H_r^\circ = \Delta H_f^\circ(\text{C}(\text{s, diamond})) - \Delta H_f^\circ(\text{C}(\text{s, graphite})) = +1.895 - 0 = +1.895 \text{ kJ/mol}$$

$$\Delta S_r^\circ = S_m^\circ(\text{C}(\text{s, diamond})) - S_m^\circ(\text{C}(\text{s, graphite})) = 2.377 - 5.740 = -3.363 \text{ J/mol}$$

$$\Delta G_r^\circ = \Delta H_r^\circ - T \Delta S_r^\circ = 1.895 \frac{\text{kJ}}{\text{mol}} - 353 \text{ K} \left(-0.003363 \frac{\text{kJ}}{\text{mol}} \right) = +3.082 \frac{\text{kJ}}{\text{mol}}$$

The reaction is spontaneous if $\Delta G_r^\circ < 0 \Rightarrow \Delta H_r^\circ - T \Delta S_r^\circ < 0$
 $\Rightarrow \Delta H_r^\circ < T \Delta S_r^\circ \Rightarrow T < \frac{\Delta H_r^\circ}{\Delta S_r^\circ}$ b/c ΔS_r° is negative.

$$\Rightarrow T < \frac{1.895 \text{ kJ/mol}}{-0.003363 \frac{\text{kJ}}{\text{mol}}} = -563 \text{ K}$$

which does NOT exist b/c the lowest temperature possible is $T^\circ = 0 \text{ K}$.

\Rightarrow So, this reaction is not spontaneous at all temperatures.

The Familiar Process of Ice Melting and Water Boiling

1. Looking back at the worksheet done in class where we calculated ΔS_{sys} , ΔS_{surr} , and $\Delta S_{\text{universe}}$ of 36 g of ice melting and in the section of this worksheet where we calculated ΔS_{sys} , ΔS_{surr} , and $\Delta S_{\text{universe}}$ of 54 g of water boiling, what was the purpose of these calculations?

To calculate $\Delta S_{\text{universe}}$ to see if this process was spontaneous under the given conditions.

There is another method of coming to these same conclusions using the Gibb's Free Energy Equation. Using data about the enthalpy and entropy of fusion and vaporization of water and the Gibb's Free Energy Equation calculate, the change of Gibb's Free Energy for each process. You should reach the same conclusions as those found on the previous worksheets.

2. Are the conditions needed in order to use the Gibb's Free Energy Equation met?

- To use $\Delta G = \Delta H - T\Delta S$ pressure and temperature need to be constant
- These conditions are met for these situations
 $P = P_{\text{atm}}$ and is constant and $T_{\text{room}} = \text{constant}$ at either 25°C, 0.1°C, 105°C or 95°C

3. What is ΔG for 36 g of ice in a room with a constant temperature of 25 °C? What can you conclude about the spontaneity of the process? $25^\circ\text{C} + 273 = 298\text{K}$

$$\Delta G = \Delta H - T\Delta S = \Delta H_{\text{fus}} \times n - T\Delta S_{\text{fus}} \times n$$

$$\Delta G = 6.01 \frac{\text{kJ}}{\text{mol}} \times 2 \text{ mol} - 298 \text{ K} \left(0.022 \frac{\text{kJ}}{\text{Kmol}} \right) (2 \text{ mol}) = -1.092 \text{ kJ}$$

The process is spontaneous b/c ΔG is negative

4. What is ΔG for 36 g of ice in a room with a constant temperature of 0.1 °C? What can you conclude about the spontaneity of the process? $0.1^\circ\text{C} + 273 = 273.1\text{K}$

$$\Delta G = 6.01 \frac{\text{kJ}}{\text{mol}} \times 2 \text{ mol} - 273.1 \text{ K} \left(0.022 \frac{\text{kJ}}{\text{Kmol}} \right) (2 \text{ mol}) = 0.0036 \text{ kJ}$$

The process is NOT spontaneous b/c ΔG is positive

5. What is ΔG for 54 g of water in an oven with a constant temperature of 105 °C? What can you conclude about the spontaneity of the process? $105^\circ\text{C} + 273 = 378\text{K}$

$$\Delta G = \Delta H - T\Delta S = \Delta H_{\text{vap}} \times n - T\Delta S_{\text{vap}} \times n$$

$$\Delta G = 40.7 \frac{\text{kJ}}{\text{mol}} \times 3 \text{ mol} - 378 \text{ K} \left(0.109 \frac{\text{kJ}}{\text{Kmol}} \right) (3 \text{ mol}) = -1.506 \text{ kJ}$$

The process is spontaneous b/c ΔG is negative

6. What is ΔG for 54 g of water in an oven with a constant temperature of 95 °C? What can you conclude about the spontaneity of the process? $95^\circ\text{C} + 273 = 368\text{K}$

$$\Delta G = 40.7 \frac{\text{kJ}}{\text{mol}} \times 3 \text{ mol} - 368 \text{ K} \left(0.109 \frac{\text{kJ}}{\text{Kmol}} \right) (3 \text{ mol}) = 1.764 \text{ kJ}$$

The process is NOT spontaneous b/c ΔG is positive

7. Which method, calculating $\Delta S_{\text{universe}}$ or calculating ΔG , do you prefer? Why?

$$\frac{36\text{g}}{18 \frac{\text{g}}{\text{mol}}} = 2 \text{ mol of } \text{H}_2\text{O}$$

For H_2O ,

$$\Delta H_{\text{fus}} = 6.01 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta S_{\text{fus}} = 22.0 \frac{\text{J}}{\text{Kmol}}$$

For H_2O ,

$$\Delta H_{\text{vap}} = 40.7 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta S_{\text{vap}} = 109 \frac{\text{J}}{\text{Kmol}}$$

$$\frac{54\text{g}}{18 \frac{\text{g}}{\text{mol}}} = 3 \text{ mol of } \text{H}_2\text{O}$$