

Thermo Test Review ! (Some data is given at the end)

1. State the first law of thermodynamics.

- Energy is conserved
- Though energy can change forms, the total amount of energy remains constant.

2. a. State the Second Law of Thermo in word form.

b. How does $\Delta G^\circ_{\text{rxn}}$ relate to the entropy change of the universe?

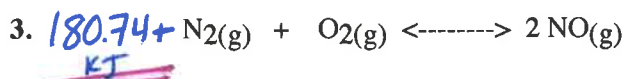
a) • The entropy of the universe must increase* for a change to occur spontaneously.

* the entropy change of the universe is equal to the entropy change of the system plus the entropy change of the surroundings.

• No process can occur that decreases the entropy of the universe unless it is accompanied by a process that increases the entropy of the universe by a larger amount.

b) $\Delta G^\circ_{\text{rxn}}$ has the opposite sign of $\Delta S_{\text{universe}}$.
so, for a rxn to be spontaneous, this means that $\Delta G^\circ_{\text{rxn}}$ must be negative, so that $\Delta S_{\text{universe}}$ is positive.

$$\left(\begin{array}{l} \Delta S_{\text{universe}} = \Delta S_{\text{rxn (system)}} + \Delta S_{\text{surroundings}} \\ \Delta S_{\text{universe}} = \Delta S_{\text{rxn}} + \left(\frac{-\Delta H_{\text{rxn}}}{T} \right) \\ -T \Delta S_{\text{universe}} = -T \Delta S_{\text{rxn}} + \Delta H_{\text{rxn}} = \Delta G_{\text{rxn}} \\ -T \Delta S_{\text{rxn}} = \Delta G_{\text{rxn}} \end{array} \right)$$



$\Delta H^\circ_{\text{rxn}} = 180.74 \text{ kJ/mole}$

$\Delta S^\circ_{\text{rxn}} = 24.7 \text{ J/mole-K}$

- a. Write the heat term into the reaction on the correct side. Is energy absorbed or released by the rxn? absorbed.
- b. Calculate $\Delta G^\circ_{\text{rxn}}$ at 25°C
- c. Is the rxn spontaneous at 25°C ? NO
- d. For what temperature range is the reaction spontaneous?
- e. Calculate K_p at 7000. Kelvin.
- f. Which K_p value make sense for this reaction at 8000. Kelvin: 1.29 or 0.775? Explain.
- g. Use the K_p value at 8000. Kelvin to calculate the $\Delta G^\circ_{\text{rxn}}$ at this temperature.

$$(b) \Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = 180.74 \frac{\text{kJ}}{\text{mole}} - (298.15 \text{ K})(.0247 \frac{\text{kJ}}{\text{mol}\cdot\text{K}})$$

$$= \boxed{173.38 \text{ kJ/mole}}$$

(d) It'll be spont. iff ΔG° is negative.

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = 0$$

$$180.74 - T(.0247) = 0$$

$$T = \frac{180.74}{.0247} = 7317 \text{ K} \rightarrow 7320 \text{ K.}$$

It'll be spontaneous
above 7320 K.

(e) first find ΔG at 7000 K

$$\Delta G = 180.74 \frac{\text{kJ}}{\text{mol}} - (7000. \text{K})(.0247 \frac{\text{kJ}}{\text{mol}\cdot\text{K}})$$

$$\Delta G = 7.84 \text{ kJ/mole}$$

$$K_p = e^{-\Delta G^\circ / RT} = e^{\frac{-7.84 \text{ kJ/mole}}{(.008314 \frac{\text{kJ}}{\text{mol}\cdot\text{K}})(7000. \text{K})}} = e^{-0.1347}$$

$$\boxed{K_p = 0.874}$$

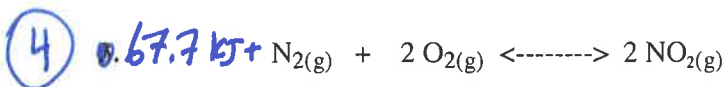
(f) 1.29 It should be

spont. at 8000 K since $8000 \text{ K} > 7320 \text{ K}$.

so $K_p > 1$, so it must be 1.29, not 0.775.

$$(g) \Delta G^\circ = -RT \ln K = -(8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}})(8000. \text{K}) \ln(1.29)$$

$$= -16937 \text{ J/mole} \text{ or } -16.937 \text{ kJ/mole} \rightarrow \boxed{-16.9 \frac{\text{kJ}}{\text{mole}}}$$



$\Delta H^\circ_{\text{rxn}} = 67.7 \text{ kJ/mole}$

$\Delta S^\circ_{\text{rxn}} = -121 \text{ J/mole}\cdot\text{K}$

- a. Write the heat term into the reaction on the correct side. Is energy absorbed or released by the rxn?
- b. Determine the sign of $\Delta S^\circ_{\text{rxn}}$. negative since 3 moles gas \rightarrow 2 moles gas ($\Delta n_{\text{gas}} = -1$)
- c. Calculate $\Delta G^\circ_{\text{rxn}}$ at 25°C
- d. Is the rxn spontaneous at 25°C ? no with fewer molecules present, entropy will be less (fewer ways to arrange the energy)
- e. For what temperature range is the reaction spontaneous?
- f. Explain your answer to (e) mathematically.
- g. Explain your answer to (e) conceptually.

c) ~~$\Delta H^\circ = 21$~~ $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = (67.7 \frac{\text{kJ}}{\text{mole}}) - (298 \text{ K})(-0.121 \frac{\text{kJ}}{\text{mol}\cdot\text{K}})$
 $= 103.758 \rightarrow 103.8 \text{ kJ/mol}$

e) It is not spontaneous for any temperature.

since ΔH is positive and ΔS is negative, no temperature could cause a negative ΔG° .

(f) $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$
 $\begin{matrix} \uparrow & & \uparrow \\ \text{positive} & & \text{negative} \\ & \underbrace{\hspace{2cm}} & \\ & \text{positive} & \end{matrix}$

so ΔG° has to be positive for any positive Kelvin Temp. so for any temp.

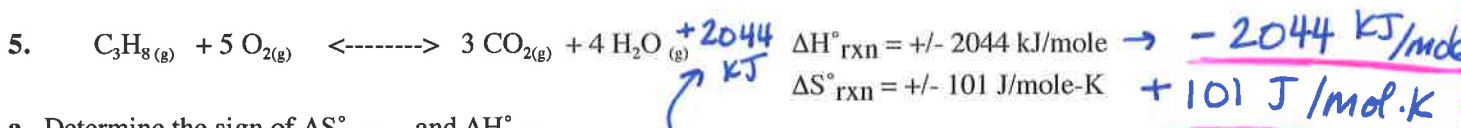
(or solve and get that ΔG° is only negative if T is below -559.5 K which is not possible)

(g) ΔH° is positive so the rxn is endothermic. This is not favorable; the rxn will need to absorb heat from the surroundings (the rest of the universe) so $\Delta S_{\text{surroundings}}$ will be negative.

ΔS_{rxn} is negative, so the molecules decrease entropy during the rxn. this is also unfavorable (the system)

$\Delta S_{\text{universe}} = \Delta S_{\text{system (rxn)}} + \Delta S_{\text{surroundings}}$

both are negative so ΔS_{univ} is negative, so not spontaneous.



- a. Determine the sign of ΔS°_{rxn} and ΔH°_{rxn} .
- b. Write the heat term into the reaction on the correct side. Is energy absorbed or released by the rxn?
- c. Calculate ΔG°_{rxn} at 25°C
- d. Is the rxn spontaneous at 25°C ? **yes**
- e. For what temperature range is the reaction spontaneous? **All temps**
- f. Explain your answer to (e) mathematically.
- g. Explain your answer to (e) conceptually.
- h. When propane gas is allowed to mix with oxygen gas at 25°C , no reaction is observed. Does this mean that $K_{eq} < 1$? Explain.
- i. What will be the enthalpy change if 10.0 grams of oxygen react?
- j. How much energy will be released, if 10.0 grams of oxygen react, at constant pressure?
- k. What mass of propane must combust in order to release 645 kJ of energy?
- l. A calorimetry experiment is done to determine ΔH of this reaction.

(c) $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = -2044 - (298.15)(0.101)$
 $= -2074.11$

$-2074 \frac{\text{kJ}}{\text{mole}}$

5.55 grams of propane are combusted while in thermal contact with 2220 grams of water. The energy from the reaction causes the water to increase in temperature from 18.4°C to 45.8°C . Calculate ΔH_{rxn} based on the data including the sign. Use kJ for the energy unit. (Ignore the heat capacity of the propane in this problem - we don't know how much its temp changed anyway)

a) ΔS must be \oplus since the # of moles of gas increases from 6 moles gas to 7 moles gas.
 ΔH must be \ominus since it shows combustion of C_3H_8 .
 Combustion is exothermic.

(f) $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

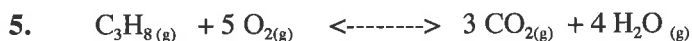
Since ΔH is negative and ΔS is positive, and since T must be positive, the equation has to yield a negative result for ΔG for any temp, so will be spontaneous at any temp.

(g) Both ΔH and ΔS contribute to the thermodynamic favorability / spontaneity of the rxn. Since ΔH is negative (exothermic rxn), the rxn will release heat into its surroundings, causing the surroundings to increase in entropy. Since ΔS is positive, the reacting molecules themselves (the system) are increasing entropy.

so $\Delta S_{universe}$ must be positive at any temp, since

$\Delta S_{universe} = \Delta S_{system} + \Delta S_{surroundings}$

Cont'd! and these are both positive.



$\Delta H^\circ_{rxn} = +/- 2044 \text{ kJ/mole}$

$\Delta S^\circ_{rxn} = +/- 101 \text{ J/mole-K}$

on previous page

- a. Determine the sign of ΔS°_{rxn} and ΔH°_{rxn} .
- b. Write the heat term into the reaction on the correct side. Is energy absorbed or released by the rxn?
- c. Calculate ΔG°_{rxn} at 25°C
- d. Is the rxn spontaneous at 25°C? _____
- e. For what temperature range is the reaction spontaneous?
- f. Explain your answer to (e) mathematically.
- g. Explain your answer to (e) conceptually.
- h. When propane gas is allowed to mix with oxygen gas at 25°C, no reaction is observed. Does this mean that $K_{eq} < 1$? Explain.
- i. What will be the enthalpy change if 10.0 grams of oxygen react?
- j. How much energy will be released, if 10.0 grams of oxygen react, at constant pressure?
- k. What mass of propane must combust in order to release 645 kJ of energy?
- l. A calorimetry experiment is done to determine ΔH of this reaction.

5.55 grams of propane are combusted while in thermal contact with 2220 grams of water. The energy from the reaction causes the water to increase in temperature from 18.4 °C to 45.8 °C. Calculate ΔH_{rxn} based on the data including the sign. Use kJ for the energy unit. (Ignore the heat capacity of the propane in this problem – we don't know how much its temp changed anyway)

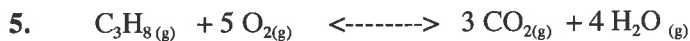
(h) **no!** $K_{eq} > 1$ at 25°C; the rxn is spontaneous at this temp acc to part (d). But the K_{eq} doesn't tell you anything about the rate of rxn; we don't know how long it will take to reach eqm. Since no rxn is observed, the rxn must be under "kinetic control." ~~meaning~~ Meaning, the activation energy must be large enough that essentially none of the collisions involve enough kinetic energy to break bonds, so the rate of rxn is negligible.

(i) $(10.0 \text{ g } O_2) \left(\frac{1 \text{ mole}}{31.9988 \text{ g}} \right) \left(\frac{-2044 \text{ kJ}}{5 \text{ moles } O_2} \right) = -127.75 \rightarrow \boxed{-128 \text{ kJ}}$
the enthalpy change

(j) 128 kJ are released. (same math as (i) except a positive amt is released.)

(k) $(645 \text{ kJ}) \left(\frac{1 \text{ mole } C_3H_8}{2044 \text{ kJ}} \right) \left(\frac{44.0962 \text{ g}}{1 \text{ mole}} \right) = \boxed{13.9 \text{ g } C_3H_8}$

(l) is on the next page.



$$\Delta H^\circ_{\text{rxn}} = \pm 2044 \text{ kJ/mole}$$

$$\Delta S^\circ_{\text{rxn}} = \pm 101 \text{ J/mole-K}$$

- Determine the sign of $\Delta S^\circ_{\text{rxn}}$ and $\Delta H^\circ_{\text{rxn}}$.
- Write the heat term into the reaction on the correct side. Is energy absorbed or released by the rxn?
- Calculate $\Delta G^\circ_{\text{rxn}}$ at 25°C
- Is the rxn spontaneous at 25°C ? _____
- For what temperature range is the reaction spontaneous?
- Explain your answer to (e) mathematically.
- Explain your answer to (e) conceptually.
- When propane gas is allowed to mix with oxygen gas at 25°C , no reaction is observed. Does this mean that $K_{\text{eq}} < 1$? Explain.
- What will be the enthalpy change if 10.0 grams of oxygen react?
- How much energy will be released, if 10.0 grams of oxygen react, at constant pressure?
- What mass of propane must combust in order to release 645 kJ of energy?
- A calorimetry experiment is done to determine ΔH of this reaction.
 - 5.55 grams of propane are combusted while in thermal contact with 2220 grams of water.
 - The energy from the reaction causes the water to increase in temperature from 18.4°C to 45.8°C .
 - Calculate ΔH_{rxn} based on the data including the sign. Use kJ for the energy unit.
 - (Ignore the heat capacity of the propane in this problem – we don't know how much its temp changed anyway)

$$(l) \quad q_{\text{H}_2\text{O}} = m_{\text{H}_2\text{O}} C_{\text{H}_2\text{O}} \Delta T_{\text{H}_2\text{O}} \quad \Delta T = 45.8 - 18.4 = 27.4^\circ\text{C}$$

$$= (2220 \text{ g}) (1 \text{ cal/g}^\circ\text{C}) (27.4^\circ\text{C}) = 60828 \text{ cal}$$

$$(60828 \text{ cal}) \left(\frac{4.184 \text{ J}}{1 \text{ cal}} \right) \left(\frac{1 \text{ kJ}}{1000 \text{ J}} \right) = 254.50 \text{ kJ}$$

absorbed by the H₂O.

$$(5.55 \text{ g C}_3\text{H}_8) \left(\frac{1 \text{ mole}}{44.0962 \text{ g}} \right) = 0.12586 \text{ moles}$$

$$\Delta H_{\text{rxn}} = \frac{-q_{\text{H}_2\text{O bath}}}{\text{moles C}_3\text{H}_8} = \frac{-254.50 \text{ kJ}}{0.12586 \text{ mole}} = -2022 \text{ kJ/mole}$$

$$\Delta H = -2020 \text{ kJ/mole}$$

3 SF

6.	ΔH°_f (kJ/mole)	ΔG°_f (kJ/mole)	S°_f (J/K-mole)
$\text{PCl}_3(\text{g})$	-288	-270.	312
$\text{PCl}_3(\text{l})$	-320.	-272	217

Thermodynamic data for $\text{PCl}_3(\text{g})$ and $\text{PCl}_3(\text{l})$ are shown above.
Use the data to estimate the normal boiling point of PCl_3 .

The normal boiling point would be the temperature at which the vapor pressure of PCl_3 is 1.00 atm



so we need to find the temp at which K_p of this rxn = 1, which would also be the temp at which $\Delta G^\circ = 0$

$$\Delta H^\circ_{\text{rxn}} = \Delta H^\circ_{\text{vap}} = (-288 \text{ kJ}) - (-320. \text{ kJ}) = 32 \frac{\text{kJ}}{\text{mole}}$$

$$\Delta S^\circ_{\text{rxn}} = \Delta S^\circ_{\text{vap}} = (312 \text{ J/K}) - (217 \text{ J/K}) = 95 \text{ J/K.mole}$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = 0$$

$$32 - T(.095 \frac{\text{kJ}}{\text{mol}\cdot\text{K}}) = 0$$

$$T = 336.84 \text{ Kelvin}$$

$$336.84 - 273.15 = 63.69 \text{ }^\circ\text{C}$$

$$\boxed{64 \text{ }^\circ\text{C}} \quad (\text{really } 15 \text{ F tho})$$

7a. Formation of ATP from ADP is not a "spontaneous" reaction.

So, how can we generate ATP in our body? What reaction do we "couple" together with the formation of ATP? Write the balanced equation for this reaction. See below.

b. What is the sign of ΔG° for each of these processes:

- (1) 1 protein \rightarrow individual amino acids \ominus (spont.)
(2) individual amino acids \rightarrow 1 protein \oplus (not spont)
(3) a bunch of glucose molecules \rightarrow a "strand" of glycogen polymer \oplus not spont
(4) 85 glucose + 510 O₂ \rightarrow 510 H₂O + 510 CO₂ \ominus Spont (C.R.)

c. Which of the above reactions are "nonspontaneous?" (2) and (3)

Since they are nonspontaneous, how do they happen in your body?

(a) To form ATP (from ADP + phosphate), we have to do

C.R. (cellular respiration): $C_6H_{12}O_6 + 6 O_2 \rightarrow 6 CO_2 + 6 H_2O$.

CR is spontaneous at all temps.

If we do C.R. together with forming (regenerating) ATP, the overall process can occur spontaneously.

(c) They can be "coupled" with the spontaneous Rxn:

$ATP \rightarrow ADP + \text{phosphate}$.

So the overall process is spontaneous.

8

a.	$2 C_{(s)} + O_{2(g)} \rightarrow 2 CO_{(g)}$	
ΔH°_f	0	0
ΔS°_f	???	205
ΔG°_f	0	0
		-110.5 kJ/mole
		197.9 J/mole K
		-137.2 kJ/mole

a. Calculate K_p at $25^\circ C$

b. Without using numbers, predict the sign of ΔS_{rxn} .

c. Determine ΔS_{rxn} .

d. Determine $\Delta S_{formation}$ of Carbon

e. For what temperature range is the rxn spontaneous?

positive since Δn_{gas} is +1

(a) $\Delta G^\circ_{rxn} = 2(-137.2) - 2(0) - 1(0) = -274.4 \text{ kJ/mole}$

(b) $K_p = e^{-\Delta G^\circ/RT} = e^{\frac{274.4 \text{ kJ/mole}}{8.314 \times 10^{-3} \frac{\text{kJ}}{\text{mol}\cdot\text{K}} \cdot 298 \text{ K}}} = e^{110.753}$

$K_p = 1.26 \times 10^{48}$

(c) $\Delta H^\circ_{rxn} = 2(-110.5) - 2(0) - 0 = -221.0 \text{ kJ/mole}$

$\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ$

$-274.4 = -221.0 - (298) \Delta S^\circ_{rxn}$

$\Delta S^\circ_{rxn} = 0.17919 \text{ kJ/mol}\cdot\text{K}$

$\Delta S^\circ_{rxn} = 179 \text{ J/K}$

(d) $\Delta S^\circ_{rxn} = 2 S_f(CO) - 2 S_f(C) - 1 S_f(O_2)$

$179.19 = 2(197.9) - 2 S_f C - 1(205)$

$S_f C = 5.805 \rightarrow 5.8 \text{ J/mol}\cdot\text{K}$

e) Spont for all temps

since ΔS°_{rxn} is \oplus and ΔH°_{rxn} is \ominus

(Combustion of propene)



a. Determine ΔH_{rxn} using heats of formation:

b. Write the heat term into the reaction on the correct side. Is energy absorbed or released by this reaction? (exo)

c. If 100. grams of propene combust, how much energy will be absorbed or released?

d. Determine the heat of combustion of propene in kJ/gram.

e. Determine ΔH_{rxn} using bond enthalpies.

$$\begin{aligned} \text{a) } \Delta H_{\text{rxn}} &= 6 \Delta H_f(\text{CO}_2) + 6 \Delta H_f(\text{H}_2\text{O}_{(\text{g})}) - 2 \Delta H_f(\text{C}_3\text{H}_6) - 9 \Delta H_f(\text{O}_2) \\ &= 6(-393.5) + 6(-241.8) - 2(20.0) - 9(0) \end{aligned}$$

$$\Delta H_{\text{rxn}} = -3851.8 \text{ kJ/mole}$$

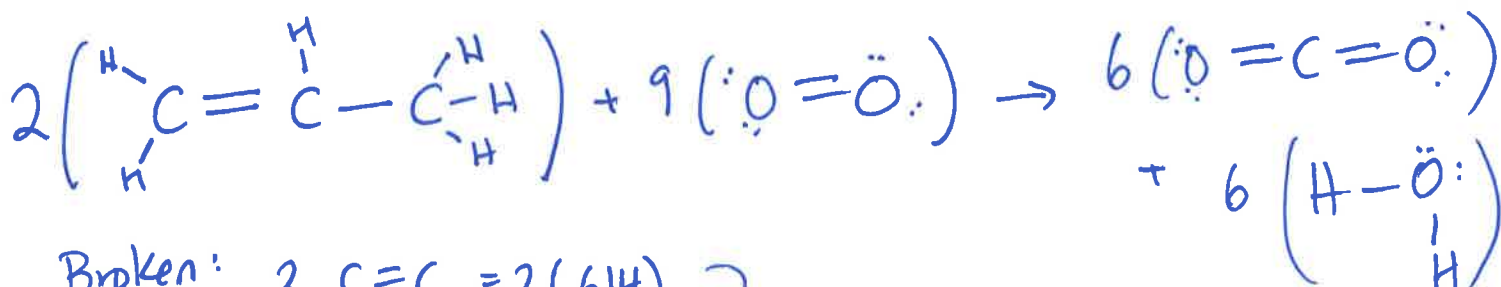
$$\text{c) } (100. \text{ g C}_3\text{H}_6) \left(\frac{1 \text{ mole}}{42.0804 \text{ g}} \right) \left(\frac{3851.8 \text{ kJ released}}{2 \text{ mole C}_3\text{H}_6} \right) = 4576.7 \text{ J}$$

4580 kJ
released

$$\text{d) } \left(\frac{3851.8 \text{ kJ}}{2 \text{ mole C}_3\text{H}_6} \right) \left(\frac{1 \text{ mole}}{42.0804 \text{ g}} \right) = 45.767 \text{ kJ/gram}$$

(or -45.767 kJ/g)

$$\text{e) } \Delta H_{\text{rxn}} = \sum \Delta H_{\text{bonds broken}} - \sum \Delta H_{\text{bonds formed}}$$



Broken:

$$\begin{aligned} 2 \text{ C}=\text{C} &= 2(614) \\ 1 \text{ C}-\text{C} &= 2(348) \\ 12 \text{ C}-\text{H} &= 12(413) \\ 9 \text{ O}=\text{O} &= 9(495) \end{aligned}$$

11335 kJ required (absorbed)
to break bonds

Formed:

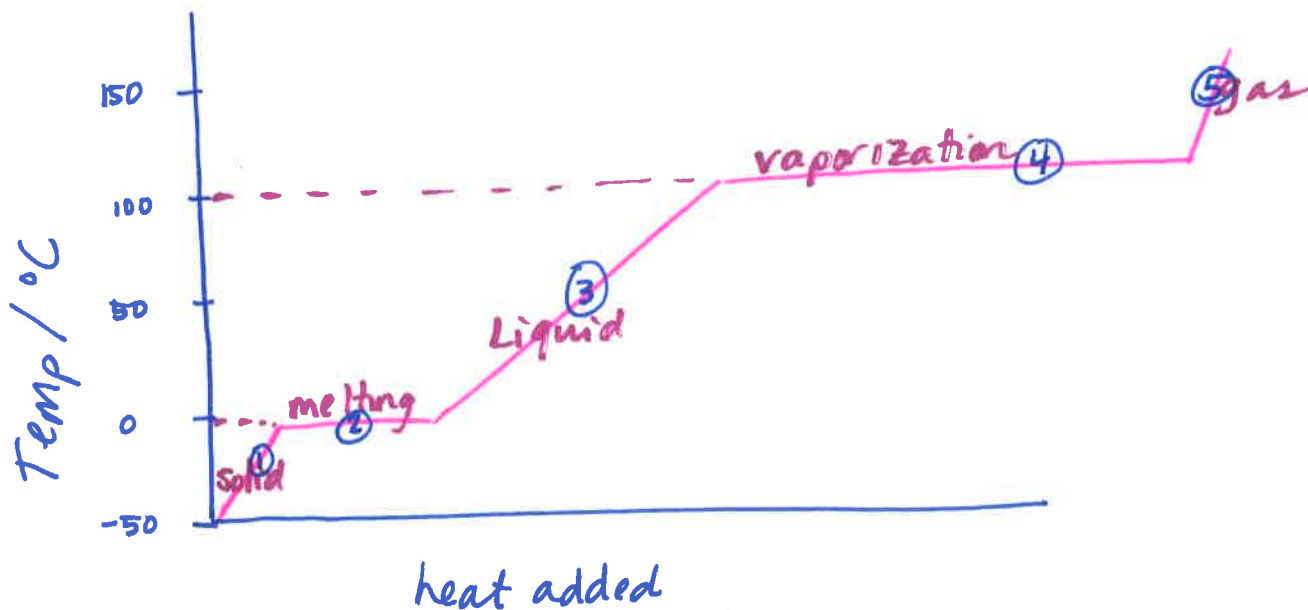
$$\begin{aligned} 12 \text{ C}=\text{O} &= 12(799) \\ 12 \text{ H}-\text{O} &= 12(463) \end{aligned}$$

15144 kJ released in forming bonds

$$\Delta H = 11335 - 15144 = -3809 \text{ kJ}$$

or kJ/mole

10. a. Sketch a graph showing temperature vs. heat added for water going from -50.0°C to 150.0°C .
 b. Calculate the total energy required for part (a) if the mass is 10.0 grams.



$$\textcircled{1} q = mc\Delta T = (10.0\text{g})(0.50 \frac{\text{cal}}{\text{g}^{\circ}\text{C}})(50.0^{\circ}\text{C}) = \underline{250} \text{ cal}$$

$$\textcircled{2} q = m\Delta H_{\text{fusion}} = (10.0\text{g})(79.9 \text{ cal/g}) = \underline{799} \text{ cal}$$

$$\textcircled{3} q = mc\Delta T = (10.0\text{g})(1 \text{ cal/g}^{\circ}\text{C})(100.0^{\circ}\text{C}) = \underline{1000} \text{ cal}$$

$$\textcircled{4} q = m\Delta H_{\text{vap}} = (10.0\text{g})(540. \text{ cal/g}) = \underline{5400} \text{ cal}$$

$$\textcircled{5} q = mc\Delta T = (10.0\text{g})(0.48 \text{ cal/g}^{\circ}\text{C})(50.0^{\circ}\text{C}) = \underline{240} \text{ cal}$$

Total: 7689 \rightarrow 7690 calories required.

11. 50.0 grams of metal at 100.0°C are placed into 125 mL of water at 4.0°C , in a "perfect calorimeter." At thermal equilibrium, the metal/water mixture is at a temperature of 6.8°C . Calculate the specific heat of the metal.

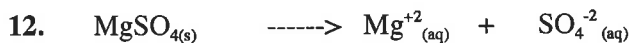
$$\left| \text{energy lost by metal} \right| = \left| \text{energy gained by H}_2\text{O} \right|$$

$$\left| mc\Delta T \right|_{\text{metal}} = \left| mc\Delta T \right|_{\text{H}_2\text{O}}$$

$$(50.0\text{g})c(93.2^{\circ}\text{C}) = (125\text{g})(1 \text{ cal/g}^{\circ}\text{C})(2.8^{\circ}\text{C})$$

\uparrow $100.0 - 6.8$ \uparrow $6.8 - 4.0$

$$c = 0.075 \text{ cal/g}^{\circ}\text{C}$$



An experiment is done to determine ΔH for the above "reaction."

87.3 grams of water and 6.48 grams of magnesium sulfate are each at a temperature of 21.4°C.

The water is placed into a calorimeter. The magnesium sulfate is added to the water, and the mixture is stirred until the MgSO_4 completely dissolves. As the salt dissolves, the solution reaches a maximum temperature of 33.3°C.

$$33.3^\circ\text{C} - 21.4^\circ\text{C} = 11.9^\circ\text{C} = \Delta T$$

Calculate ΔH for the "reaction" (so, calculate " $\Delta H_{\text{solution}}$ of MgSO_4 ") in kJ/mole.

Assume that the calorimeter contents have a specific heat of 4.184 J/g°C.

$$q = m c \Delta T = (87.3\text{g} + 6.48\text{g})(4.184\text{J/g}^\circ\text{C})(11.9^\circ\text{C}) = 4669.3\text{J}$$

$$(4669.3\text{J}) \left(\frac{1\text{kJ}}{1000\text{J}} \right) = 4.6693\text{kJ absorbed by H}_2\text{O etc and released by the "rxn"}$$

$$(6.48\text{g MgSO}_4) \left(\frac{1\text{mole}}{120.3666\text{g}} \right) = 0.053836\text{ moles MgSO}_4$$

$$|\Delta H_{\text{rxn}}| = \frac{4.6693\text{kJ}}{0.053836\text{ moles}} = 86.733 \rightarrow 86.7\text{ kJ/mole}$$

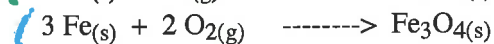
but since the H_2O (etc) temp went up, the "rxn" must have released energy into the H_2O , so the rxn is exothermic, so ΔH is negative.

$$\Delta H = -86.7 \frac{\text{kJ}}{\text{mole}}$$

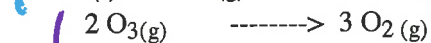
13. Given the following:



$$\Delta H_{\text{rxn}} = -964.96\text{ kJ/mole}$$



$$\Delta H_{\text{rxn}} = -1117.1\text{ kJ/mole}$$



$$\Delta H_{\text{rxn}} = -284\text{ kJ/mole}$$

Use Hess's Law to calculate ΔH_{rxn} for: $6\text{Fe}_2\text{O}_3(s) \rightarrow 4\text{Fe}_3\text{O}_4(s) + \text{O}_2(g)$



$$\Delta H = 469.36\text{ kJ} \rightarrow 469\text{ kJ or kJ/mole}$$

DATA!!!

Bond Enthalpies: (kJ per mole of bonds broken)

Cl-Cl: 242	H-O: 463	C-O: 358	C=O: 799	O-O: 146
C-H: 413	C-C: 348	C=C: 614	C-C: 839	O=O: 495

Enthalpy of formation: (kJ per mole)

$\text{C}_3\text{H}_6(g)$ 20.0	$\text{CO}_2(g)$ -393.5	$\text{H}_2\text{O}(l)$ -285.8	$\text{H}_2\text{O}(g)$ -241.8	$\text{CaCl}_2(s)$ -
791.4				

specific heat (cal/g°C)	$\text{H}_2\text{O}(s)$ 0.50	$\text{H}_2\text{O}(l)$ 1.00	$\text{H}_2\text{O}(g)$ 0.48
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heat of fusion for water: 79.9 cal/g heat of vaporization for water: 540. cal/g

(#14 is on the next page)

14. a. For each of the following phase changes, state whether the ΔH of the phase change is positive or negative, and whether the ΔS of the phase change is positive or negative.

	ΔH	ΔS	
melting	<u>+</u>	<u>+</u>	Solid \rightarrow liquid
freezing	<u>-</u>	<u>-</u>	liquid \rightarrow solid
condensation	<u>-</u>	<u>-</u>	gas \rightarrow liquid <u>or</u> gas \rightarrow solid
vaporization	<u>+</u>	<u>+</u>	liquid \rightarrow gas
sublimation	<u>+</u>	<u>+</u>	Solid \rightarrow gas

b. Which of the phase changes are more "favorable" at high temp than at low temp, and how does this relate to the sign of the ΔH and ΔS ?

Melting, Vaporization, and sublimation are all thermodynamically favorable at high temp, but not at low T. These phase changes have a positive ΔH and ΔS .

$$\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ$$

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As Temperature (T) increases, ΔG becomes more negative, so favorability increases.

ΔG must be negative for process to be spontaneous/favored.

c. Which of the phase changes are more "favorable" at low temp than at high temp, and how does this relate to the sign of the ΔH and ΔS ?

Freezing and Condensation are thermodynamically favorable at low temp, but not at high temp.

$$\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ$$

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As temp increases, the ΔG becomes more positive, so favorability decreases.

(ΔG° will be positive at high T \rightarrow not favorable.)
 (ΔG° will be negative at low T \rightarrow favorable.)