

This print-out should have 30 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

LDE Thermodynamic Theory 003

001 6.0 points

Which of the following statements concerning the first law of thermodynamics is/are true?

- I) The internal energy of the universe is always increasing.
- II) Internal energy lost by a system is always gained by the surroundings.
- III) The universe is an isolated system.

- 1. I and III only
- 2. I only
- 3. II only
- 4. I and II only
- 5. I, II and III
- 6. II and III only **correct**
- 7. III only

Explanation:

Statement I is false; the first law states that the energy of the universe is conserved, in other words a constant value. Statement II and III are true; internal energy in the universe is conserved, and thus energy lost by the system is always gained by the surroundings. The universe is the most obvious example of an isolated system in that energy and matter are conserved in the universe.

LDE Definitions in Thermo 001

002 6.0 points

Which of the following is **not** a definition of enthalpy or change in enthalpy?

- 1. A measure of a system's ability to change the entropy of its surroundings.
- 2. A correction for internal energy that accounts for pressure-volume work.

3. A measure of the motional energy of a system. **correct**

4. A measure of the heat of a system at constant pressure.

Explanation:

Total motional energy is a a measure of the internal energy (U , or sometimes E) of a system, not the enthalpy (H).

LDE Sign Conventions 002

003 6.0 points

You observe that carbon dioxide sublimes. Which of the following statements about the signs of this process is/are true?

- I) Work (w) is positive.
- II) Heat (q) is negative.
- III) Change in Gibbs free energy (ΔG) is positive.
- IV) Change in entropy (ΔS) is positive.

- 1. III and IV
- 2. I and II
- 3. II and III
- 4. I only
- 5. I, II and III
- 6. IV only **correct**

Explanation:

Sublimation results in a significant increase in the volume of the system, allowing it to do work on its surroundings, i.e. the pressure-volume work function is negative. Sublimation is also an endothermic process, making heat positive. Since the process described "happens" as a given in the problem, the change in free energy must be negative. Change in entropy must be positive since a solid is becoming a gas.

LDE Definitions in Thermo 002

004 6.0 points

Which of the following quantities is **not** path

independent?

1. pressure (P)
2. entropy (S)
3. enthalpy (H)
4. heat (q) **correct**
5. volume (V)

Explanation:

Heat quantitatively describes a transition **between** states, but not a state itself, and is thus a process quantity (path function), not a state function.

LDE Definitions in Thermo 003

005 6.0 points

Which of the following reactions **is** an enthalpy of formation reaction?

1. $2\text{Fe}(\text{s}) + 3/2\text{O}_2(\text{g}) \longleftrightarrow \text{Fe}_2\text{O}_3(\text{s})$ **correct**
2. $\text{Hg}(\text{s}) + 1/2\text{O}_2(\text{g}) \longleftrightarrow \text{HgO}(\text{s})$
3. $\text{C}_{\text{diamond}}(\text{s}) + 2\text{H}_2(\text{g}) \longleftrightarrow \text{CH}_4(\text{g})$
4. $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \longleftrightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}$

Explanation:

Formation reactions describe production of exactly one mole of one product from stoichiometric quantities of elements in their standard states.

LDE Definitions in Thermo 004

006 6.0 points

Which of the following would experience the smallest increase in temperature if 1 kJ of heat were added to it?

1. 1 g of water
2. 1 g of copper metal
3. 10 g of copper metal

4. 10 g of water **correct**

Explanation:

Heat capacity is an extensive property, that reflects the amount of energy required to raise the temperature of an object. Because water has a much higher heat capacity than metals (to some extent a consequence of IMF), and because 10 grams is more than 1 gram, the 10 gram sample of water would experience a much smaller increase in temperature than the other choices.

LDE Bomb Calorimeter 002

007 6.0 points

If we set up a bomb calorimetry experiment to determine the molar enthalpy of combustion of ethene (C_2H_4) using 1 L of water as our heat sink, 2.805 g of ethene, and measure an initial and final temperature of 25.20°C and 58.92°C , respectively, what will be the experimentally determined molar enthalpy of combustion of ethene? Assume the density of water is $1.00\text{ g} \cdot \text{mL}^{-1}$. Assume the calorimeter itself absorbs no heat. The specific heat capacity of water is $4.184\text{ J} \cdot \text{g}^{-1} \cdot \text{K}^{-1}$.

1. $-14.11\text{ kJ} \cdot \text{mol}^{-1}$
2. $-141.1\text{ kJ} \cdot \text{mol}^{-1}$
3. $-1,411\text{ kJ} \cdot \text{mol}^{-1}$ **correct**
4. $-141,100\text{ kJ} \cdot \text{mol}^{-1}$
5. $-14,110\text{ kJ} \cdot \text{mol}^{-1}$

Explanation:

$$\Delta T = T_f - T_i = 58.92^\circ\text{C} - 25.20^\circ\text{C} = 33.72^\circ\text{C} = 33.72\text{ K}$$

$$m = 1\text{ L} \cdot \frac{1000\text{ mL}}{\text{L}} \cdot \frac{1.00\text{ g}}{\text{mL}} = 1000\text{ g}$$

$$n = 2.805\text{ g ethene} \cdot \frac{28.05\text{ g}}{\text{mol}} = 0.1\text{ mol}$$

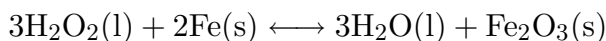
$$\Delta H_{rxn} = -\Delta H_{cal} = -mc\Delta T$$

$$\frac{-mc\Delta T}{n} = -\frac{1000 \text{ g} \cdot \frac{4.184 \text{ J}}{\text{g}\cdot\text{K}} \cdot 33.72 \text{ K}}{0.1 \text{ mol}}$$

$$= -1,411 \text{ kJ} \cdot \text{mol}^{-1}$$

LDE Hess' Law 002
008 6.0 points

Based on enthalpy of formation data, calculate ΔH for the reaction:



1. $-992 \text{ kJ} \cdot \text{mol}^{-1}$
2. $992 \text{ kJ} \cdot \text{mol}^{-1}$
3. $1,118 \text{ kJ} \cdot \text{mol}^{-1}$
4. $-1,118 \text{ kJ} \cdot \text{mol}^{-1}$ **correct**

Explanation:

$$\Delta H_{rxn} = \Sigma \Delta H_{f,products} + \Sigma \Delta H_{f,reactants}$$

$$= (3 \cdot \Delta H_{f,\text{H}_2\text{O}(\text{l})} + \Delta H_{f,\text{Fe}_2\text{O}_3(\text{s})}) - (3 \cdot \Delta H_{f,\text{H}_2\text{O}_2(\text{l})} + 2 \cdot \Delta H_{f,\text{Fe}(\text{s})})$$

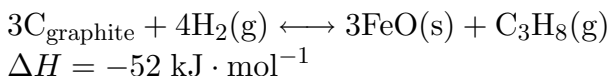
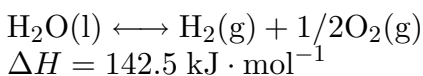
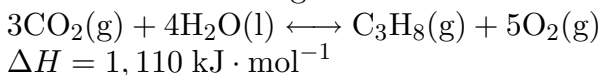
$$=$$

$$(3 \cdot -285.83 \text{ kJ} \cdot \text{mol}^{-1} + -824.2 \text{ kJ} \cdot \text{mol}^{-1}) - (3 \cdot -187.78 \text{ kJ} \cdot \text{mol}^{-1} + 2 \cdot 0)$$

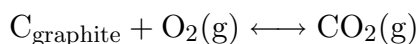
$$= -1,118 \text{ kJ} \cdot \text{mol}^{-1}$$

LDE Hess' Law 003
009 6.0 points

Given the following data:



calculate ΔH for the reaction



1. $543 \text{ kJ} \cdot \text{mol}^{-1}$
2. $-592 \text{ kJ} \cdot \text{mol}^{-1}$
3. $-543 \text{ kJ} \cdot \text{mol}^{-1}$
4. $1202 \text{ kJ} \cdot \text{mol}^{-1}$
5. $-197 \text{ kJ} \cdot \text{mol}^{-1}$ **correct**

Explanation:

$$\Delta H_{rxn} = \frac{11}{3}(1,110 \text{ kJ} \cdot \text{mol}^{-1})$$

$$- \frac{4}{3}(142.5 \text{ kJ} \cdot \text{mol}^{-1}) - \frac{1}{3}(-52 \text{ kJ} \cdot \text{mol}^{-1})$$

$$= -197 \text{ kJ} \cdot \text{mol}^{-1}$$

LDE Internal Energy Calc 002
010 6.0 points

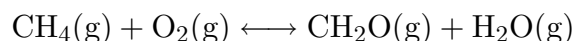
How much energy would be associated with the translational motion of 0.5 moles of sugar, $\text{C}_6\text{H}_{12}\text{O}_6$?

1. $6kT$
2. $0.75kT$
3. $6RT$
4. $18kT$
5. $0.75RT$ **correct**
6. $18RT$

Explanation:

LDE Bond Enthalpy 002
011 6.0 points

Using the provided bond enthalpy data, calculate the change in enthalpy for the following reaction:



1. $577 \text{ kJ} \cdot \text{mol}^{-1}$

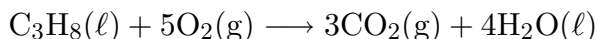
2. $-710 \text{ kJ} \cdot \text{mol}^{-1}$
3. $-349 \text{ kJ} \cdot \text{mol}^{-1}$ **correct**
4. $349 \text{ kJ} \cdot \text{mol}^{-1}$
5. $710 \text{ kJ} \cdot \text{mol}^{-1}$
6. $-577 \text{ kJ} \cdot \text{mol}^{-1}$

Explanation:

$$\begin{aligned}\Delta H_{rxn} &= \sum BE_{\text{reactants}} - \sum BE_{\text{products}} \\ &= 4 \cdot 412 \text{ kJ} \cdot \text{mol}^{-1} + 496 \text{ kJ} \cdot \text{mol}^{-1} \\ &\quad - 2 \cdot 412 \text{ kJ} \cdot \text{mol}^{-1} - 743 \text{ kJ} \cdot \text{mol}^{-1} \\ &\quad - 2 \cdot 463 \text{ kJ} \cdot \text{mol}^{-1} \\ &= -349 \text{ kJ} \cdot \text{mol}^{-1}\end{aligned}$$

LDE Thermodynamic Work 002**012** 6.0 points

Approximate the work for the following reaction at 300 K.



1. 2.5 kJ
2. -2.5 kJ
3. 0 kJ
4. 5 kJ **correct**
5. -5 kJ

Explanation:

At room temperature (300 K), the product of the RT term in $W = -\Delta n_{\text{gas}}RT$ is approximately 2.5 kJ. Since Δn_{gas} is -2 for this reaction, the work function is approximately 5 kJ.

LDE Definitions in Thermo 005**013** 6.0 points

Which of the following is **not** a definition of internal energy or change in internal energy?

1. The sum of heat (q) and work (w).
2. A measure of the change in heat of a system at constant volume.
3. The total energy content of a system.
4. A measure of the spontaneity of a reaction. **correct**

Explanation:

Gibbs free energy is a measure of the spontaneity of a reaction. All of the other statements are mathematical identities describing internal energy or prose restatements of the same.

LDE Bomb Calorimeter 003**014** 6.0 points

Which of the following statements concerning calorimetry is/are true?

- I) The calories listed on a food package is actually a measure of kilocalories.
- II) In a calorimeter, the heat sink should have a low heat capacity.
- III) A bomb calorimetry measurement is made at constant pressure.

1. I only **correct**
2. II only
3. II, III
4. I, II, III
5. III only
6. I, III
7. I, II

Explanation:

Statement I is true; food calories are actual thousands of calories (kilocalories). Statement II is false; water is a common heat sink and is preferred because it has a high heat capacity. Statement III is false; in a bomb calorimeter, the volume is held constant.

LDE Internal Energy Calc 003**015** 6.0 points

If an MP3 player does 200 kJ of work and releases 100 kJ of heat, what is the change in internal energy for the MP3 player?

1. –100 kJ
2. 300 kJ
3. –300 kJ **correct**
4. 100 kJ

Explanation:

$$\Delta U = q + w$$

$$= -100 \text{ kJ} + -200 \text{ kJ} = -600 \text{ kJ}$$

LDE Entropy Change Calc 002**016** 6.0 points

Rank the following reactions from least to greatest in terms of change in entropy (ΔS_{rxn}):

- a) $\text{KNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \longrightarrow \text{NaNO}_3(\text{aq}) + \text{KCl}(\text{aq})$
- b) $4\text{Ag}(\text{s}) + \text{O}_2(\text{g}) \longrightarrow 2\text{Ag}_2\text{O}(\text{s})$
- c) $2\text{H}_2\text{O}_2(\ell) \longrightarrow 2\text{H}_2\text{O}(\text{g}) + \text{O}_2(\text{g})$
- d) $\text{NaHCO}_3(\text{s}) \longrightarrow \text{NaOH}(\text{s}) + \text{CO}_2(\text{g})$

1. $d < c < a < b$
2. $a < b < d < c$
3. $b < a < d < c$ **correct**
4. $a < d < b < c$
5. $b < a < c < d$
6. $c < d < b < a$

Explanation:

Reaction b has a decrease in total moles and a decrease in moles of gas; ΔS_{rxn} is a negative number. In reaction a, no changes in phase and no changes in number of moles

occur; ΔS_{rxn} will be nearly zero. Reaction d evolves a single mole of gas; ΔS_{rxn} will be a small positive number. Reaction c evolves 3 moles of gas; ΔS_{rxn} will be a large positive number.

LDE Entropy Change Calc 003**017** 6.0 points

What is the change in entropy for the condensation of water? The standard molar entropy of $\text{H}_2\text{O}(\text{g})$ and of $\text{H}_2\text{O}(\ell)$ are $188.83 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$ and $69.91 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$, respectively.

1. $-258.74 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$
2. $258.74 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$
3. $-118.92 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$ **correct**
4. $118.92 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$

Explanation:

The change in entropy for condensation of water can be described by the equation $\text{H}_2\text{O}(\text{g}) \longleftrightarrow \text{H}_2\text{O}(\ell)$.

$$\begin{aligned} \Delta S_{\text{condensation}} &= \Sigma S_{m,\text{liquid}} - \Sigma S_{m,\text{gas}} \\ &= \\ &69.91 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} - 188.83 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \\ &= -118.92 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \end{aligned}$$

LDE Thermodynamic Theory 004**018** 6.0 points

Which of the following statements concerning the second and third laws of thermodynamics is/are true?

- I) When the change in entropy of the system is equal in magnitude and opposite in sign to the change in entropy of the surroundings, the change in entropy of the universe is zero.
- II) The change in entropy of the universe can be rewritten as $\Delta G_{\text{system}}/T$ (do you mean to have a negative sign here?)
- III) In a perfect, pure crystal at absolute zero the entropy of the system is zero.

1. II only

2. I, II

3. II, III

4. I, II, III **correct**

5. III only

6. I, III

7. I only

Explanation:

Statement I is true; the second law equation states that the change in entropy of the universe is the sum of the change in entropy of the system and the surroundings. Statement II is true; this is a mathematical identity derived from the second law equation. Statement III is true; the third law states that as the temperature of a system approaches absolute zero, the entropy of the system approaches its minimum, which in the case of a perfect crystalline solid, is zero.

LDE Stat Thermo Theory 002
019 6.0 points

When calculating the positional entropy of a system using the Boltzmann formula ($S = k \ln W$), which of the following statements is/are true?

- I) As the complexity of a system increases, W increases.
- II) As the intermolecular forces in a system increase, W increases.
- III) The absolute entropy of a system cannot be negative.

1. I, II, III

2. III only

3. II only

4. I only **correct**

5. II, III

6. I, III

7. I, II

Explanation:

Statement I is true; W is a restatement of the "complexity" of the system. Statement II is false; as IMF increases, it biases positions toward an arrangement that maximizes electrostatic interactions and thus "orders" the system decreasing its "complexity." Statement III is true; W can never be less than one, and this positional entropy can never be less than zero.

LDE Stat Thermo Entropy 002
020 6.0 points

What would the positional entropy of 2 moles of BFCl_2 be at absolute zero?

1. $9.13 \text{ J} \cdot \text{K}^{-1}$ 2. $18.27 \text{ J} \cdot \text{K}^{-1}$ **correct**3. $3.03 \times 10^{-23} \text{ J} \cdot \text{K}^{-1}$ 4. $1.52 \times 10^{-23} \text{ J} \cdot \text{K}^{-1}$ **Explanation:**

BFCl_2 would have 3 possible orientations. $S = k \ln(3^n) = 2R \cdot \ln 3 = 18.27 \text{ J} \cdot \text{K}^{-1}$

LDE Ranking Entropies 002
021 6.0 points

Rank the following compounds in terms of decreasing molar entropy: $\text{NO}(\text{s})$, $\text{NO}_2(\text{g})$, $\text{NO}_2(\ell)$, $\text{NO}(\ell)$.

1. $\text{NO}_2(\text{g}) > \text{NO}_2(\ell) > \text{NO}(\ell) > \text{NO}(\text{s})$
correct

2. $\text{NO}(\text{s}) > \text{NO}_2(\ell) > \text{NO}(\ell) > \text{NO}_2(\text{g})$ 3. $\text{NO}(\text{s}) > \text{NO}_2(\text{g}) > \text{NO}_2(\ell) > \text{NO}(\ell)$ 4. $\text{NO}_2(\text{g}) > \text{NO}(\text{s}) > \text{NO}(\ell) > \text{NO}_2(\ell)$ 5. $\text{NO}_2(\text{g}) > \text{NO}(\ell) > \text{NO}_2(\ell) > \text{NO}(\text{s})$ **Explanation:**

$\text{NO}_2(\text{g})$ and $\text{NO}(\text{s})$ obviously belong on the extremes, being the most and least entropic

on account of their phase. The two liquids can be compared based on their number of bonds, and so $\text{NO}_2(\ell)$ has greater molar entropy than $\text{NO}(\ell)$.

LDE Ideality Ranking 001

022 6.0 points

Assuming you have one mole of each of the following, rank them in terms of decreasing value of the term W in the Boltzmann equation: HF, NaF, ClF, F_2 .

1. HF > NaF > F_2 > ClF
2. F_2 > ClF > HF > NaF **correct**
3. ClF > HF > NaF > F_2
4. NaF > F_2 > ClF > HF

Explanation:

Since the strength of the IMF determines the extent of the perturbation of W from its ideal value, this problem essentially requires ranking by increasing IMF. The greater the IMF, the closer the value of W will be to unity.

LDE Entropy Change Calc 004

023 6.0 points

If a system at 273°C absorbs 1000 J of heat, what is its change in entropy?

1. $1.83 \text{ J} \cdot \text{K}^{-1}$ **correct**
2. $-3.66 \text{ J} \cdot \text{K}^{-1}$
3. $-1.83 \text{ J} \cdot \text{K}^{-1}$
4. $3.66 \text{ J} \cdot \text{K}^{-1}$

Explanation:

Heat absorbed by a system will increase its entropy by an amount equal to q/T , in this case $1000\text{J}/546\text{K} = 1.83 \text{ J} \cdot \text{K}^{-1}$.

LDE Thermo 2nd Law Calc 002

024 6.0 points

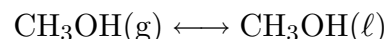
Calculate $\Delta S_{\text{universe}}$ for the condensa-

tion of methanol (CH_3OH) at 10°C if the S_m° for $\text{CH}_3\text{OH}(\text{g})$ and $\text{CH}_3\text{OH}(\ell)$ are $240 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$ and $127 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$, respectively. Assume the standard boiling point of methanol is 65°C .

1. $-113 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$
2. $135 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$
3. $-135 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$
4. $22 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$ **correct**
5. $113 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$
6. $-22 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$

Explanation:

Since the system in this case is condensing methanol,



$$\begin{aligned} \Delta S_{\text{system}} &= \Sigma S_{m,\text{products}} - \Sigma S_{m,\text{reactants}} \\ &= 127 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} - 240 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \\ &= -113 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \end{aligned}$$

$$\begin{aligned} \Delta H_{\text{system}} &= T \Delta S_{\text{system}} \\ &= 338 \text{ K} \cdot -113 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \\ &= -38,194 \text{ J} \cdot \text{mol}^{-1} \end{aligned}$$

Since the only process during a phase change is heat exchange,

$$\begin{aligned} \Delta S_{\text{surroundings}} &= -\Delta H_{\text{system}}/T_{\text{surroundings}} \\ &= 38,194 \text{ J} \cdot \text{mol}^{-1} / 283 \text{ K} \\ &= 134.96 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \end{aligned}$$

$$\begin{aligned} \Delta S_{\text{universe}} &= \Delta S_{\text{system}} + \Delta S_{\text{surroundings}} \\ &= -113 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} + 135 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \\ &= 22 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \end{aligned}$$

LDE Thermo 2nd Law Calc 003

025 6.0 points

If a given process increases the entropy of the universe by $10,000 \text{ J} \cdot \text{K}^{-1}$ in spite of the the entropy of the system decreasing by $1,000 \text{ J} \cdot \text{K}^{-1}$, what must the change in entropy of the surroundings have been?

1. $9,000 \text{ J} \cdot \text{K}^{-1}$

2. $-9,000 \text{ J} \cdot \text{K}^{-1}$

3. $11,000 \text{ J} \cdot \text{K}^{-1}$ **correct**

4. $-11,000 \text{ J} \cdot \text{K}^{-1}$

Explanation:

$$\begin{aligned}\Delta S_{\text{universe}} &= \Delta S_{\text{system}} + \Delta S_{\text{surroundings}} \\ 10,000 \text{ J} \cdot \text{K}^{-1} & \\ &= -1,000 \text{ J} \cdot \text{K}^{-1} + \Delta S_{\text{surroundings}} \\ \Delta S_{\text{surroundings}} &= 11,000 \text{ J} \cdot \text{K}^{-1}\end{aligned}$$

LDE Gibbs Eqn 001**026** 6.0 points

Phosphine (the common name for PH₃, a highly toxic gas used for fumigation), has a $\Delta H_{\text{vap}}^{\circ} = 14.6 \text{ kJ} \cdot \text{mol}^{-1}$ and a $S_{\text{vap}}^{\circ} = 78.83 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$. What is the normal boiling point of phosphine expressed in centigrade?

1. $-87.8 \text{ }^{\circ}\text{C}$ **correct**

2. $185.2 \text{ }^{\circ}\text{C}$

3. $273 \text{ }^{\circ}\text{C}$

4. $-0.2 \text{ }^{\circ}\text{C}$

Explanation:

Because boiling is an equilibrium process, $\Delta G_{\text{vap}}^{\circ} = 0 = \Delta H_{\text{vap}}^{\circ} - T\Delta S_{\text{vap}}^{\circ}$.

$$\begin{aligned}\text{And so } T\Delta S_{\text{vap}}^{\circ} &= \Delta H_{\text{vap}}^{\circ} \\ \text{and } T &= \Delta H_{\text{vap}}^{\circ} / \Delta S_{\text{vap}}^{\circ} \\ &= 14,600 \text{ J} \cdot \text{mol}^{-1} / 78.83 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \\ &= 185.2 \text{ K} \\ &= -87.8 \text{ }^{\circ}\text{C}\end{aligned}$$

LDE Gibbs Eqn 002**027** 6.0 points

For a given reaction, if $\Delta H_{\text{rxn}}^{\circ}$ is (negative/positive/either) and $\Delta S_{\text{rxn}}^{\circ}$ is (negative/positive/either), then the value of $\Delta G_{\text{rxn}}^{\circ}$ will always decrease as you raise the temperature.

1. either, negative

2. positive, either

3. negative, either

4. positive, negative

5. either, positive **correct**

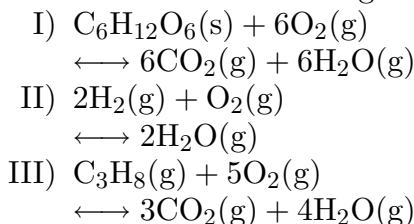
6. negative, positive

Explanation:

Only the sign of $\Delta S_{\text{rxn}}^{\circ}$ determines how $\Delta G_{\text{rxn}}^{\circ}$ will be effected by changes in temperature. When $\Delta S_{\text{rxn}}^{\circ}$ is positive, $\Delta G_{\text{rxn}}^{\circ}$ will always decrease in value as you raise the temperature. This can be intuited from the Maxwell equation, $\Delta G = \Delta H - T\Delta S$. Note that only ΔS is multiplied by T , not ΔH .

LDE Reaction Spontaneity 002**028** 6.0 points

Which of the following combustion reactions does not occur at high temperature?



1. II only **correct**

2. I only

3. I, III

4. III only

5. II, III

6. I, II

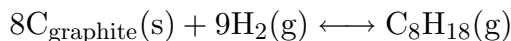
7. I, II, III

Explanation:

Only reaction II results in a decrease in entropy, and is thus becomes less spontaneous and eventually non-spontaneous as temperature increases.

LDE Gibbs Eqn 003**029** 6.0 points

Octane is C_8H_{18} . Consider the reaction below:



$$\Delta G_{rxn}^{\circ} = 17.32 \text{ kJ} \cdot \text{mol}^{-1}$$

The (formation/decomposition) of octane is (spontaneous/non-spontaneous) and therefore octane is (stable/unstable).

1. formation, spontaneous, stable
2. decomposition, spontaneous, stable
3. decomposition, non-spontaneous, stable
4. decomposition, non-spontaneous, unstable
5. formation, non-spontaneous, unstable **correct**

Explanation:

The reaction written above describes the formation of octane and has a positive change in Gibbs free energy, i.e. is non-spontaneous, meaning octane must be thermodynamically unstable.

Explanation:

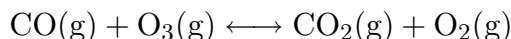
$$\begin{aligned} \Delta H_{rxn} &= \Sigma H_{f,products} - \Sigma H_{f,reactants} \\ &= (-393.5 \text{ kJ} \cdot \text{mol}^{-1}) \\ &\quad - (-110.5 \text{ kJ} \cdot \text{mol}^{-1} + 143 \text{ kJ} \cdot \text{mol}^{-1}) \\ &= -426 \text{ kJ} \cdot \text{mol}^{-1} \end{aligned}$$

$$\begin{aligned} \Delta S_{rxn} &= \Sigma S_{m,products} - \Sigma S_{m,reactants} \\ &= (214 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} + 205 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}) \\ &\quad - (198 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} + 239 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}) \\ &= -18 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \\ &= -0.018 \text{ kJ} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \end{aligned}$$

$$\begin{aligned} \Delta G_{rxn}^{\circ} &= \Delta H_{rxn}^{\circ} - T \Delta S_{rxn}^{\circ} \\ &= -426 \text{ kJ} \cdot \text{mol}^{-1} - \\ &\quad (273 \text{ K})(-0.018 \text{ kJ} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}) \\ &= -421 \text{ kJ} \cdot \text{mol}^{-1} \end{aligned}$$

LDE Gibbs Eqn 004**030** 6.0 points

Consider the reaction below:



Using the provided table values, calculate ΔG_{rxn}° if it is performed under standard conditions.

1. $-4,488 \text{ kJ} \cdot \text{mol}^{-1}$
2. $421 \text{ kJ} \cdot \text{mol}^{-1}$
3. $-421 \text{ kJ} \cdot \text{mol}^{-1}$ **correct**
4. $4,488 \text{ kJ} \cdot \text{mol}^{-1}$