

This print-out should have 6 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

LDE Thermodynamic Theory 010

001 10.0 points

Which law of thermodynamics governs the spontaneity of reactions?

1. The 5th Law
2. The 3rd Law
3. The 0th Law
4. The 1st Law
5. The 2nd Law **correct**

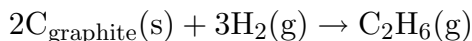
Explanation:

The second law of thermodynamics states that the entropy of the universe is always increasing. Consequently, only processes which increase the overall entropy of the universe satisfy the second law and happen spontaneously.

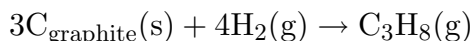
LDE Gibbs Stability Ranking 001

002 10.0 points

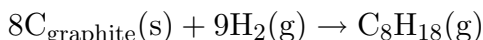
Consider the formation reactions below and pick the most stable species from the answer choices.



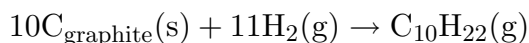
$$\Delta G_f^\circ = -7.86 \text{ kcal} \cdot \text{mol}^{-1}$$



$$\Delta G_f^\circ = -5.614 \text{ kcal} \cdot \text{mol}^{-1}$$



$$\Delta G_f^\circ = 4.14 \text{ kcal} \cdot \text{mol}^{-1}$$



$$\Delta G_f^\circ = 8.23 \text{ kcal} \cdot \text{mol}^{-1}$$

1. $\text{C}_3\text{H}_8(\text{g})$
2. $\text{C}_2\text{H}_6(\text{g})$ **correct**
3. $\text{C}_{10}\text{H}_{22}(\text{g})$
4. $\text{C}_8\text{H}_{18}(\text{g})$

Explanation:

The formation of ethane is the most exergonic of the formation reactions and thus ethane is the most stable of the species formed.

LDE Entropy Change Calc 008

003 10.0 points

Calculate the change in entropy of the surroundings (ΔS_{surr}) if a system absorbs 5.79 kJ of heat and the surroundings are at 95.7 °C.

1. 60.5 J · K⁻¹
2. -60.5 J · K⁻¹
3. -15.7 J · K⁻¹ **correct**
4. 15.7 J · K⁻¹

Explanation:

$$\Delta S_{\text{surr}} = -\frac{\Delta H_{\text{sys}}}{T_{\text{surr}}} = \frac{-5,790 \text{ J}}{368.7 \text{ K}} = -15.7 \text{ J} \cdot \text{K}^{-1}$$

LDE Temperature and Phase Changes 003

004 10.0 points

Based on the enthalpy of sublimation ($\Delta H_{\text{sub}} = 393.5 \text{ kJ} \cdot \text{mol}^{-1}$) and entropy of sublimation ($\Delta S_{\text{sub}} = 2.023 \text{ kJ} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$) of carbon dioxide, at what temperature does this phase transition occur?

1. 0.2 °C
2. 78.5 °C
3. -78.5 K

4. $-78.5\text{ }^\circ\text{C}$ correct

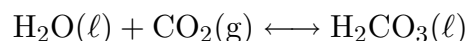
5. 0.2 K

Explanation:

$$T_{sub} = \frac{\Delta H_{sub}}{\Delta S_{sub}} = \frac{393.5\text{ kJ} \cdot \text{mol}^{-1}}{2.023\text{ kJ} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}} = 194.5\text{ K} = -78.5\text{ }^\circ\text{C}$$

LDE Gibbs Eqn 009
005 10.0 points

Consider the reaction below:



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

	ΔH_f° ($\text{kJ} \cdot \text{mol}^{-1}$)	S_m° ($\text{J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$)
$\text{CO}_2(\text{g})$	-393.51	213.74
$\text{H}_2\text{O}(\ell)$	-285.83	69.91
$\text{H}_2\text{CO}_3(\ell)$	-1108.51	156.41

1. $-143.23\text{ kJ} \cdot \text{mol}^{-1}$

2. $-301.93\text{ kJ} \cdot \text{mol}^{-1}$

3. $-391.25\text{ kJ} \cdot \text{mol}^{-1}$ correct

4. $301.93\text{ kJ} \cdot \text{mol}^{-1}$

5. $391.25\text{ kJ} \cdot \text{mol}^{-1}$

Explanation:

$$\begin{aligned} \Delta H_{rxn} &= \Sigma H_{f,products} - \Sigma H_{f,reactants} \\ &= (-1108.51\text{ kJ} \cdot \text{mol}^{-1}) \\ &\quad - (-285.83\text{ kJ} \cdot \text{mol}^{-1} - 393.51\text{ kJ} \cdot \text{mol}^{-1}) \\ &= -429.17\text{ kJ} \cdot \text{mol}^{-1} \end{aligned}$$

$$\begin{aligned} \Delta S_{rxn} &= \Sigma S_{m,products} - \Sigma S_{m,reactants} \\ &= (156.41\text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}) \\ &\quad - (69.91\text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \\ &\quad + 213.74\text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}) \end{aligned}$$

$$\begin{aligned} &= -127.24\text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \\ &= -0.12724\text{ kJ} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \end{aligned}$$

$$\begin{aligned} \Delta G_{rxn}^\circ &= \Delta H_{rxn}^\circ - T\Delta S_{rxn}^\circ \\ &= -429.17\text{ kJ} \cdot \text{mol}^{-1} - \\ &\quad (298\text{ K})(-0.12724\text{ kJ} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}) \\ &= -391.25\text{ kJ} \cdot \text{mol}^{-1} \end{aligned}$$

LDE Thermo 2nd Law Calc 006
006 10.0 points

Your roommate left 1 kg of ice out on the counter last night and all of it melted. Given that $\Delta H = 6.01\text{ kJ} \cdot \text{mol}^{-1}$ and $\Delta S = 22.0\text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$ for water melting and that the temperature in the room was $25\text{ }^\circ\text{C}$, by how much in total has your roommate increased the entropy of the universe?

1. $0\text{ J} \cdot \text{K}^{-1}$

2. $101.6\text{ J} \cdot \text{K}^{-1}$ correct

3. $2340\text{ J} \cdot \text{K}^{-1}$

4. $1.83\text{ J} \cdot \text{K}^{-1}$

Explanation:

$$\begin{aligned} \Delta S_{surr} &= -\frac{\Delta H}{T} = -\frac{6010\text{ J} \cdot \text{mol}^{-1}}{298\text{ K}} \\ &= -20.17\text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \end{aligned}$$

$$\begin{aligned} \Delta S_{univ} &= \Delta S_{sys} + \Delta S_{surr} \\ &= 1.83\text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \end{aligned}$$

$$\begin{aligned} 1\text{ kg of water} &= \frac{1000\text{ g}}{18.0148\text{ g} \cdot \text{mol}^{-1}} \\ &= 55.5\text{ moles} \end{aligned}$$

$$\begin{aligned} \text{so the total increase is } &(55.5 \cdot 1.83)\text{ J} \cdot \text{K}^{-1} \\ &= 101.6\text{ J} \cdot \text{K}^{-1} \end{aligned}$$