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 5^{th} Law

2. The 3^{rd} Law

3. The 0^{th} Law

4. The 1^{st} Law

5. The 2^{nd} Law **correct**

Explanation:
The second law of thermodynamics states that the entropy of the universe is always increasing. Consequently, only processes which increases 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.

$$2\mathrm{C}_{\text{graphite}}(\mathrm{s}) + 3\mathrm{H}_2(\mathrm{g}) \to \mathrm{C}_2\mathrm{H}_6(\mathrm{g})$$

$$\Delta \mathrm{G}^\circ_\mathrm{f} = -7.86 \text{ kcal} \cdot \text{mol}^{-1}$$

$$3\mathrm{C}_{\text{graphite}}(\mathrm{s}) + 4\mathrm{H}_2(\mathrm{g}) \to \mathrm{C}_3\mathrm{H}_8(\mathrm{g})$$

$$\Delta \mathrm{G}^\circ_\mathrm{f} = -5.614 \text{ kcal} \cdot \text{mol}^{-1}$$

$$8\mathrm{C}_{\text{graphite}}(\mathrm{s}) + 9\mathrm{H}_2(\mathrm{g}) \to \mathrm{C}_8\mathrm{H}_{18}(\mathrm{g})$$

$$\Delta \mathrm{G}^\circ_\mathrm{f} = 4.14 \text{ kcal} \cdot \text{mol}^{-1}$$

$$10\mathrm{C}_{\text{graphite}}(\mathrm{s}) + 11\mathrm{H}_2(\mathrm{g}) \to \mathrm{C}_{10}\mathrm{H}_{22}(\mathrm{g})$$

$$\Delta \mathrm{G}^\circ_\mathrm{f} = 8.23 \text{ kcal} \cdot \text{mol}^{-1}$$

1. C_3H_8(g)

2. C_2H_6(g) **correct**

3. $C_{10}H_{22}$(g)

4. C_8H_{18}(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. $60.5 \text{ J} \cdot \text{K}^{-1}$

2. $-60.5 \text{ J} \cdot \text{K}^{-1}$

3. $-15.7 \text{ J} \cdot \text{K}^{-1}$ **correct**

4. $15.7 \text{ J} \cdot \text{K}^{-1}$

Explanation:

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

LDE Temperature and Phase Changes 003
004 10.0 points

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

1. 0.2 °C

2. 78.5 °C

3. −78.5 K

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

5. 0.2 K

Explanation:
$T_{sub} = \dfrac{\Delta H_{sub}}{\Delta S_{sub}} = \dfrac{393.5\ \mathrm{kJ\cdot mol^{-1}}}{2.023\ \mathrm{kJ\cdot mol^{-1}\cdot K^{-1}}} =$
$194.5\ \mathrm{K} = -78.5\ ^\circ\mathrm{C}$

LDE Gibbs Eqn 009
005 10.0 points

Consider the reaction below:

$$H_2O(\ell) + CO_2(g) \longleftrightarrow H_2CO_3(\ell)$$

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

	ΔH°_f ($\mathrm{kJ\cdot mol^{-1}}$)	S°_m ($\mathrm{J\cdot mol^{-1}\cdot K^{-1}}$)
$CO_2(g)$	-393.51	213.74
$H_2O(\ell)$	-285.83	69.91
$H_2CO_3(\ell)$	-1108.51	156.41

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

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

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

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

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

Explanation:
$\Delta H_{rxn} = \Sigma H_{f,products} - \Sigma H_{f,reactants}$
$= (-1108.51\ \mathrm{kJ\cdot mol^{-1}})$
$-(-285.83\ \mathrm{kJ\cdot mol^{-1}} - 393.51\ \mathrm{kJ\cdot mol^{-1}})$
$= -429.17\ \mathrm{kJ\cdot mol^{-1}}$

$\Delta S_{rxn} = \Sigma S_{m,products} - \Sigma S_{m,reactants}$
$= (156.41\ \mathrm{J\cdot mol^{-1}\cdot K^{-1}})$
$-(69.91\ \mathrm{J\cdot mol^{-1}\cdot K^{-1}}$
$+213.74\ \mathrm{J\cdot mol^{-1}\cdot K^{-1}})$
$= -127.24\ \mathrm{J\cdot mol^{-1}\cdot K^{-1}}$
$= -0.12724\ \mathrm{kJ\cdot mol^{-1}\cdot K^{-1}}$

$\Delta G^\circ_{rxn} = \Delta H^\circ_{rxn} - T\Delta S^\circ_{rxn}$
$= -429.17\ \mathrm{kJ\cdot mol^{-1}} -$
$(298\ \mathrm{K})(-0.12724\ \mathrm{kJ\cdot mol^{-1}\cdot K^{-1}})$
$= -391.25\ \mathrm{kJ\cdot mol^{-1}}$

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 $\Delta \mathrm{H} = 6.01\ \mathrm{kJ\cdot mol^{-1}}$ and $\Delta \mathrm{S} = 22.0\ \mathrm{J\cdot mol^{-1}\cdot K^{-1}}$ for water melting and that the temperature in the room was $25\ ^\circ\mathrm{C}$, by how much in total has has your roommate increased the entropy of the universe?

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

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

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

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

Explanation:
$\Delta \mathrm{S_{surr}} = -\dfrac{\Delta \mathrm{H}}{\mathrm{T}} = -\dfrac{6010\ \mathrm{J\cdot mol^{-1}}}{298\ \mathrm{K}}$

$= -20.17\ \mathrm{J\cdot mol^{-1}\cdot K^{-1}}$

$\Delta \mathrm{S_{univ}} = \Delta \mathrm{S_{sys}} + \Delta \mathrm{S_{surr}}$

$= 1.83\ \mathrm{J\cdot mol^{-1}\cdot K^{-1}}$

$1\ \mathrm{kg\ of\ water} = \dfrac{1000\ \mathrm{g}}{18.0148\ \mathrm{g\cdot mol^{-1}}}$

$= 55.5$moles

so the total increase is $(55.5 \cdot 1.83)\ \mathrm{J\cdot K^{-1}}$

$= 101.6\ \mathrm{J\cdot K^{-1}}$