This print-out should have 25 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

Please refer to the cover page for needed thermodynamic values and formulas.

001 10.0 points
An isolated system allows for the flow of...?

1. none of these correct
2. sound waves
3. kinetic energy
4. matter
5. heat

Explanation:
In the natural sciences an isolated system is a physical system without any external exchange - neither matter nor energy (as heat or work) can enter or exit, but can only move around inside.

002 10.0 points
Which of the following is true of a general thermodynamic state function?

1. The change in the value of a state function is always negative for a spontaneous reaction.
2. The value of the state function remains constant.
3. The value of a state function does NOT change with a change in temperature of a process.
4. The change of the value of a state function is independent of the path of a process. correct
5. The change in the value of the state function is always positive for endothermic processes.

Explanation:
A change in a state function describes a difference between the two states. It is independent of the process or pathway by which the change occurs.

003 10.0 points
Consider the reaction
\[ \text{C}_6\text{H}_{14}(ℓ) + 9.5 \text{O}_2(g) \rightarrow 6 \text{CO}_2(g) + 7 \text{H}_2\text{O}(ℓ) \]
at constant pressure. Which response is true?

1. No work is done as the reaction occurs.
2. Work is done by the system as the reaction occurs.
3. Work is done on the system as the reaction occurs. correct
4. Work may be done on or by the system as the reaction occurs, depending upon the temperature.

Explanation:
For \( P = \text{const}, \)
\[ w = -P \Delta V = -(\Delta n) RT , \]
\[ \Delta n = n_f - n_i = -3.5 \text{ mol gas} . \]

For \( -\Delta n, \) \( w \) will be positive, which indicates that work was done on the system, which progresses as the reaction progresses.

004 10.0 points
Consider the following specific heats: copper, 0.384 J/g·°C; lead, 0.159 J/g·°C; water, 4.18 J/g·°C; glass, 0.502 J/g·°C. If the same amount of heat is added to identical masses of each of these substances, which substance attains the highest temperature? (Assume that they all have the same initial temperature.)

1. copper
2. water
3. lead correct
4. glass

**Explanation:**

005 10.0 points
Which of the following statements is/are true concerning the first law of thermodynamics?
I) The internal energy \((U)\) of the universe is conserved.
II) The internal energy of a system plus that of its surroundings is conserved.
III) The change in internal energy \((\Delta U)\) of a system and its surroundings can have the same sign.

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

**Explanation:**
Statement I and II are true; the first law states that the internal energy of the universe is conserved and since the system plus the surroundings is the universe, their sum is also conserved. Statement III is false; for example, if both the system and its surroundings had a positive change in internal energy, then the internal energy of the universe would have increased or decreased - in violation of the first law.

006 10.0 points

1.95 mol of an ideal gas at 300 K and 3.00 atm expands from 16 L to 28 L and a final pressure of 1.20 atm in two steps:
(1) the gas is cooled at constant volume until its pressure has fallen to 1.20 atm, and
(2) it is heated and allowed to expand against a constant pressure of 1.20 atm un-
til its volume reaches 28 L.
Which of the following is CORRECT?

1. \(w = 0\) for the overall process
2. \(w = -6.03\) kJ for the overall process
3. \(w = -4.57\) kJ for (1) and \(w = -1.46\) kJ for (2)
4. \(w = 0\) for (1) and \(w = -1.46\) kJ for (2) correct
5. \(w = -4.57\) kJ for the overall process

**Explanation:**
For step (1): If there is no change in volume, \(w = 0\).
For step (2): For expansion against a constant external pressure,

\[
w = -P_{\text{ext}} \Delta V = (-1.2 \text{ atm})(18 \text{ L} - 6 \text{ L}) \\
\times (101.325 \text{ J} \cdot \text{L}^{-1} \cdot \text{atm}^{-1}) = -1.45908 \text{ kJ}.
\]

The total work done by the system would be the sum of the work for each step.

007 10.0 points

A CD player and its battery together do 500 kJ of work, and the battery also releases 250 kJ of energy as heat and the CD player releases 50 kJ as heat due to friction from spinning. What is the change in internal energy of the system, with the system regarded as the battery and CD player together?

1. \(-800\) kJ correct
2. +200 kJ
3. \(-200\) kJ
4. \(-700\) kJ
5. \(-750\) kJ

**Explanation:**
Heat from the CD player is \(-50\) kJ.
Heat from the battery is $-500 \text{ kJ}$.
Work from both together on the surroundings is $-250 \text{ kJ}$.
This question is testing your ability to see what the system is, and then look at ONLY the energy flow for the system. Here the system is the battery and the CD player together.

$$\Delta U = q + w$$

$$= [-50 \text{ kJ} + (-250 \text{ kJ})] + (-500 \text{ kJ})$$

$$= -800 \text{ kJ}$$

**008 10.0 points**
The specific heat of liquid water is 4.184 $\text{J/g}\cdot\text{°C}$, and of steam 2.03 $\text{J/g}\cdot\text{°C}$. The heat of vaporization of water ($\ell$) is 2.26 $\text{kJ/g}$ and its boiling point is 100$\text{°C}$. What is the total heat flow when 18 grams of water at 12$\text{°C}$ are heated to become steam at 109$\text{°C}$?

1. 44.4 kJ
2. 47.6 kJ correct
3. under 28 kJ
4. 48.9 kJ
5. over 55 kJ
6. 31.7 kJ
7. 40.7 kJ

**Explanation:**

\[ \begin{align*}
18 \text{ g} & \quad \text{step 1} \quad 18 \text{ g} & \quad \text{step 2} \\
\text{H}_2\text{O}(\ell) & \quad \text{H}_2\text{O}(\ell) & \quad \text{H}_2\text{O}(\ell) \\
12\text{°C} & \quad 100\text{°C} & \quad 109\text{°C} \\
\text{step 3} & \\
\text{H}_2\text{O}(\ell) & \quad \text{H}_2\text{O}(\ell) & \quad \text{H}_2\text{O}(\ell) \\
18 \text{ g} & \quad 18 \text{ g} & \quad 18 \text{ g} \\
100\text{°C} & \quad 100\text{°C} & \quad 109\text{°C}
\end{align*} \]

Step 1: \[ \frac{4.184 \text{ J}}{\text{g} \cdot \text{°C}} \cdot (18 \text{ g}) \cdot (100 - 12)\text{°C} \]

\[ = 6,627 \text{ J} \]

Step 2: \[ \frac{2.26 \text{ kJ}}{\text{g}} \cdot (18 \text{ g}) \cdot \frac{1000 \text{ J}}{1 \text{ kJ}} \]

\[ = 40,680 \text{ J} \]

Step 3: \[ \frac{2.03 \text{ J}}{\text{g} \cdot \text{°C}} \cdot (18 \text{ g}) \cdot (109 - 100)\text{°C} \]

\[ = 329 \text{ J} \]

Total = \[ 6627 \text{ J} + 329 \text{ J} + 40,680 \text{ J} \]

\[ = 47,636 \text{ J} = 47.636 \text{ kJ} \]

**009 10.0 points**
Refer to the potential energy diagram shown below.

**Reaction progress**

What is the change in enthalpy ($\Delta H$) for the reaction \[ A \rightarrow B \]?

1. +50 kJ, endothermic correct
2. −50 kJ, exothermic
3. +50 kJ, exothermic
4. −250 kJ, endothermic
5. −150 kJ, endothermic
6. −50 kJ, endothermic
7. +300 kJ, exothermic
8. −150 kJ, exothermic
9. +300 kJ, endothermic
10. −250 kJ, exothermic

**Explanation:**

\[ \Delta H_i = \Delta H_A = 250 \text{ kJ} \]
\[ \Delta H_f = \Delta H_B = 300 \text{ kJ} \]

\[ \Delta H = \Delta H_f - \Delta H_i \]

\[ = 300 \text{ kJ} - 250 \text{ kJ} \]

\[ = 50 \text{ kJ} \]

Notice that $\Delta H$ is positive and therefore the reaction is endothermic.
Which of the following is NOT a feature of the bomb calorimetry apparatus used to measure the internal energy of a reaction?

1. The heat capacity of the calorimeter should be known to accurately correct for any heat lost to it.

2. The thermometer is inserted directly into the reaction vessel to measure $\Delta T$ of the reaction. correct

3. Large quantities of water surrounding the reaction vessel absorb the majority of the heat loss.

4. The volume of the reaction vessel is held constant to eliminate energy released as work.

5. The large heat capacity of water is beneficial in measuring heat released by combustion reactions.

Explanation:
The thermometer is placed in the water that surrounds the reaction vessel.

\[ m_{C_6H_{14}} = 1.00 \text{ g} \quad m_{\text{water}} = 1502 \text{ g} \]

\[ SH = 4.184 \frac{\text{J}}{\text{g} \cdot ^\circ \text{C}} \quad HC = 4042 \frac{\text{J}}{^\circ \text{C}} \]

\[ \Delta T = 29.30^\circ \text{C} - 22.64^\circ \text{C} = 6.66^\circ \text{C} \]

The increase in the water temperature is $29.30^\circ \text{C} - 22.64^\circ \text{C} = 6.66^\circ \text{C}$. The amount of heat responsible for this increase in temperature for 1502 g of water is

\[ q = (6.66^\circ \text{C}) \left( 4.184 \frac{\text{J}}{\text{g} \cdot ^\circ \text{C}} \right) (1502 \text{ g}) \]

\[ = 41854 \text{ J} = 41.85 \text{ kJ} \]

The amount of heat responsible for the warming of the calorimeter is

\[ q = (6.66^\circ \text{C})(4042 \frac{\text{J}}{^\circ \text{C}}) \]

\[ = 26920 \text{ J} = 26.92 \text{ kJ} \]

The amount of heat released on the reaction is thus 41.85 kJ + 26.92 kJ = 68.77 kJ per g of n-hexane. Per mol of n-hexane, this becomes

\[ \left( 68.77 \frac{\text{kJ}}{\text{g}} \right) \left( 86.1 \frac{\text{g}}{\text{mol}} \right) = 5921 \text{ kJ/mol} \]

However, since heat is released, the sign is negative.

011 10.0 points

A 1.00 g sample of n-hexane ($C_6H_{14}$) undergoes complete combustion with excess $O_2$ in a bomb calorimeter. The temperature of the 1502 g of water surrounding the bomb rises from 22.64°C to 29.30°C. The heat capacity of the hardware component of the calorimeter (everything that is not water) is 4042 J/°C. What is $\Delta U$ for the combustion of n-$C_6H_{14}$? One mole of n-$C_6H_{14}$ is 86.1 g. The specific heat of water is 4.184 J/g·°C.

1. $-1.15 \times 10^4$ kJ/mol
2. $-4.52 \times 10^3$ kJ/mol
3. $-7.40 \times 10^4$ kJ/mol
4. $-5.92 \times 10^3$ kJ/mol correct
5. $-9.96 \times 10^3$ kJ/mol

Explanation:
A formation reaction produces exactly one mole of one product from elements in their standard states.

012 10.0 points

Which of the reactions below is a formation reaction?

1. $H_2(g) + C_{\text{graphite}}(s) + \frac{3}{2}O_2(g) \rightarrow CO_2(g) + H_2O(g)$

2. $2H_2(g) + 2C_{\text{graphite}}(s) + O_2(g) \rightarrow 2CH_2O(\ell)$

3. $2Fe(s) + \frac{3}{2}O_2(g) \rightarrow Fe_2O_3(s)$ correct

4. $N_2(\ell) + 2H_2(g) \rightarrow N_2H_4(\ell)$

Explanation:
A formation reaction produces exactly one mole of one product from elements in their standard states.

013 10.0 points
Calculate the standard reaction enthalpy for the reaction \[ \text{N}_2\text{H}_4(\ell) + \text{H}_2(\text{g}) \rightarrow 2 \text{NH}_3(\text{g}) \] given \[ \text{N}_2\text{H}_4(\ell) + \text{O}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g}) \] \[ \Delta H^\circ = -543 \text{ kJ} \cdot \text{mol}^{-1} \] \[ 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g}) \] \[ \Delta H^\circ = -484 \text{ kJ} \cdot \text{mol}^{-1} \] \[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g}) \] \[ \Delta H^\circ = -92.2 \text{ kJ} \cdot \text{mol}^{-1} \]

1. $-1119 \text{ kJ} \cdot \text{mol}^{-1}$
2. $-151 \text{ kJ} \cdot \text{mol}^{-1}$ correct
3. $-935 \text{ kJ} \cdot \text{mol}^{-1}$
4. $-243 \text{ kJ} \cdot \text{mol}^{-1}$
5. $-59 \text{ kJ} \cdot \text{mol}^{-1}$

**Explanation:**
We need to reverse the second reaction and add them:

\[ \text{N}_2\text{H}_4(\ell) + \text{O}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g}) \] \[ \Delta H^\circ = -543 \text{ kJ} \cdot \text{mol}^{-1} \]

\[ 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g}) \] \[ \Delta H^\circ = -484 \text{ kJ} \cdot \text{mol}^{-1} \]

\[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g}) \] \[ \Delta H^\circ = -92.2 \text{ kJ} \cdot \text{mol}^{-1} \]

\[ \text{N}_2\text{H}_4(\ell) + \text{H}_2(\text{g}) \rightarrow 2 \text{NH}_3(\text{g}) \] \[ \Delta H = \sum E_{\text{break}} - \sum E_{\text{make}} \]

\[ = [(\text{C}≡\text{C}) + (\text{H}≡\text{Br})] \]
\[ - [(\text{C}≡\text{H}) + (\text{C}≡\text{Br}) + (\text{C}≡\text{C})] \]
\[ = (602 \text{ kJ/mol} + 366 \text{ kJ/mol}) \]
\[ - [413 \text{ kJ/mol} + 285 \text{ kJ/mol} + 346 \text{ kJ/mol}] \]
\[ = -76 \text{ kJ/mol}, \]

so $76 \text{ kJ/mol}$ of heat was released.

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Estimate the heat released when ethene (\(\text{CH}_2\equiv\text{CH}_2\)) reacts with HBr to give \(\text{CH}_3\text{CH}_2\text{Br}\).

1. $200 \text{ kJ/mol}$
2. $76 \text{ kJ/mol}$ correct
3. $1036 \text{ kJ/mol}$

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Consider a reaction that is non-spontaneous at all temperatures. What would be the signs of \(\Delta G_{\text{sys}}\), \(\Delta H_{\text{surr}}\), and \(\Delta S_{\text{univ}}\) respectively for such a reaction?

1. $+, -, +$
2. $-, +, -$ correct
3. $+, -, -$
4. $-, +, +$
5. $+, +, +$

**Explanation:**
For a reaction that is non-spontaneous at all temperatures, the free energy of the system will increase and the entropy of the universe will decrease. Such a reaction must be endothermic, and the heat it gains will be lost from the surroundings.
Which one shows the substances in the decreasing order of their molar entropy?

1. C(s), H₂O(ℓ), H₂O(g), H₂O(s)
2. H₂O(g), H₂O(ℓ), H₂O(g), C(s) correct
3. None of these
4. H₂O(s), H₂O(ℓ), H₂O(g), C(s)
5. C(s), H₂O(g), H₂O(ℓ), H₂O(s)
6. C(s), H₂O(s), H₂O(ℓ), H₂O(g)

Explanation:
Gases will have a higher entropy than liquids so we expect H₂O(ℓ) to have the lowest molar entropy. The gases will increase in entropy in the order Ne(g) < Ar(g) < CO₂(g). Ne and Ar are both atoms so they should have less entropy than a molecular substance, which has more complexity. Ar will have a higher entropy than Ne because it has a larger mass and more fundamental particles. The correct order is H₂O(ℓ) < Ne(g) < Ar(g) < CO₂(g).

4. V
5. II and III correct

Explanation:
\[ R = 8.314 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \]
Assume 1 mol in each case. Entropy decreases if \( \Delta S \) is negative.
For the oxygen gas pressure doubling isothermally, \( P_2 = 2 P_1 \) and
\[
\Delta S = n R \ln \left( \frac{P_1}{P_2} \right) \\
= (1 \text{ mol})(8.314 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}) \ln \left( \frac{1}{2} \right) \\
= -5.76 \text{ J} \cdot \text{K}^{-1}
\]
We expect a negative answer since pressure increased.
For the CO₂ gas volume expanding 10× isothermally, \( V_2 = 10 V_1 \) and
\[
\Delta S = n R \ln \left( \frac{V_2}{V_1} \right) \\
= (1.00 \text{ mol})(8.314 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}) \times \ln(10) \\
= +38.29 \text{ J} \cdot \text{K}^{-1}
\]
We expect a positive answer since volume increased.
For the nitrogen gas compressed to \( \frac{1}{2} \) original volume isothermally, \( V_1 = 2 V_2 \) and
\[
\Delta S = n R \ln \left( \frac{V_2}{V_1} \right) \\
= (1 \text{ mol})(8.314 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}) \ln \left( \frac{1}{2} \right) \\
= -5.76 \text{ J} \cdot \text{K}^{-1}
\]
We expect a negative answer since volume decreased.
For the cooling glass of water,
\[ T = 30^\circ \text{C} + 273.15 = 303.15 \text{ K} \]
\[
\Delta S = \frac{q}{T} = \frac{-200 \text{ J}}{303.15 \text{ K}} = -0.6597 \text{ J} \cdot \text{K}^{-1}
\]
The last situation (heating the 1 mol of He) does not give enough data to calculate an answer but from the formula

\[ \Delta S = n C_{p,m} \ln \left( \frac{T_2}{T_1} \right) \]

\( n = 1 \) mol and for a monoatomic ideal gas \( C_{p,m} = 2.5 \, R \). Finally if the temperature increases this means \( T_2 > T_1 \) so \( \ln \left( \frac{T_2}{T_1} \right) \) will be positive. We expect a positive answer since temperature increased.

**018 10.0 points**
What is the entropy at \( T = 0 \) K for one mole of chloroform (CHCl\(_3\))?  

1. \( 1.9 \times 10^{-23} \) J K\(^{-1}\)  
2. 1.38 J K\(^{-1}\)  
3. 11.5 J K\(^{-1}\) **correct**  
4. –11.5 J K\(^{-1}\)  
5. 0 J K\(^{-1}\)  

**Explanation:** Since the question is concerned with the residual entropy of a mole of chloroform, which as 4 orientations, we can use the Boltzmann equation to calculate the residual entropy.  

\[ S = N_a k \ln W = R \ln 4 = 11.5 \text{ J K}^{-1} \]

**019 10.0 points**
What is the total non-vibrational internal energy of 10 nitrous oxide (N\(_2\)O) molecules?  

1. 15 k T  
2. 10 k T  
3. 25 k T **correct**  
4. 10 R T  
5. 25 R T  

**Explanation:** If you burn gasoline in an engine to move a car, you are ultimately converting chemical potential energy into kinetic energy. But much of this energy is lost as heat. There is NO way to make any energy conversion 100% efficient.

**020 10.0 points**
Advertising claims sometimes state that adding something mechanical to a car’s engine will allow it to recover 100% of the energy that comes from burning gasoline. You should be skeptical of such claims because they violate the  

1. first law of thermodynamics.  
2. activation energy requirements of all chemical reactions.  
3. law of conservation of matter.  
4. second law of thermodynamics. **correct**  

**Explanation:** If you burn gasoline in an engine to move a car, you are ultimately converting chemical potential energy into kinetic energy. But much of this energy is lost as heat. There is NO way to make any energy conversion 100% efficient.

**021 10.0 points**
Calculate \( \Delta S_{\text{surr}}^o \) at 298 K for the reaction  

\[ 6 \text{C(s)} + 3 \text{H}_2(\text{g}) \rightarrow \text{C}_6\text{H}_6(\ell) \]

\( \Delta H_r^o = +49.0 \text{ kJ\cdotmol}^{-1} \) and \( \Delta S_r^o = -253 \text{ J\cdotK}^{-1}\cdot\text{mol}^{-1}. \)

1. \(-417 \text{ J\cdotK}^{-1}\cdot\text{mol}^{-1}\)  
2. \(-164 \text{ J\cdotK}^{-1}\cdot\text{mol}^{-1} \) **correct**  
3. \(+253 \text{ J\cdotK}^{-1}\cdot\text{mol}^{-1}\)
4. \(-253 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}\)

5. \(+164 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}\)

**Explanation:**

\[ \Delta H^\circ_r = 49000 \text{ J} \cdot \text{mol}^{-1} \]

\[ T = 298 \text{ K} \]

\[ \Delta S^\circ_{\text{surr}} = \frac{q_{\text{surr}}}{T} = \frac{-q}{T} = \frac{-\Delta H^\circ_r}{298 \text{ K}} \]

\[ = \frac{-(+49000 \text{ J} \cdot \text{mol}^{-1})}{298 \text{ K}} \]

\[ = -164.43 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}. \]

**022 10.0 points**

For the evaporation of water from an open pan at 25°C, the values of \(\Delta S\) for the water, the surroundings, and the universe must be, respectively,

1. positive, positive, positive.
2. positive, negative, zero.
3. None of these is correct.
4. negative, negative, negative.
5. positive, negative, positive. **correct**

**Explanation:**

The process is spontaneous, which means \(\Delta S_{\text{universe}} > 0\) according to the Second Law of Thermodynamics.

Entropy \((S)\) is high for systems with high degrees of freedom, disorder or randomness and low for systems with low degrees of freedom, disorder or randomness.

\[ S(\text{g}) > S(\ell) > S(\text{s}). \]

The system is

\[ \text{H}_2\text{O}(\ell) \rightarrow \text{H}_2\text{O}(\text{g}) \]

so \(\Delta S_{\text{system}} > 0\).

The entropy of the surroundings must be negative. Energy is removed from the surroundings to get the water to evaporate.

**023 10.0 points**

For the reaction

\[ 2 \text{SO}_3(\text{g}) \rightarrow 2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \]

\[ \Delta H^\circ_r = +198 \text{ kJ} \cdot \text{mol}^{-1} \text{ at } 298 \text{ K.} \]

Which statement is true for this reaction?

1. The reaction is driven by the enthalpy.
2. The reaction will not be spontaneous at any temperature.
3. The reaction will not be spontaneous at high temperatures.
4. \(\Delta G^\circ_r\) will be positive at high temperatures.
5. \(\Delta G^\circ_r\) will be negative at high temperatures. **correct**

**Explanation:**

\(\Delta G = \Delta H - T\Delta S\) is used to predict spontaneity. (\(\Delta G\) is negative for a spontaneous reaction.) \(\Delta H\) is positive and \(T\) is always positive. For the reaction 2 mol gas \(\rightarrow\) 3 mol gas. The more moles of gas, the higher the disorder, so \(\Delta S\) is positive and \(\Delta G = (+) - T(+).\) For \(\Delta G\) to be negative, \(T\) must be large.

**024 10.0 points**

What is \(\Delta G^\circ_r\) for the combustion of liquid \(n\)-pentane?

1. 3389 kJ/mol
2. \(-383\) kJ/mol
3. 383 kJ/mol
4. \(-3389\) kJ/mol **correct**
5. \(-451\) kJ/mol
6. 451 kJ/mol

**Explanation:**

The combustion of liquid \(n\)-pentane oc-
curs according to the following reaction.
\[ C_5H_{12}(ℓ) + 8 O_2(g) \rightarrow 5 CO_2(g) + 6 H_2O(ℓ) \]

\[ \Delta H = \left[ 5(-394) + 6(-286) \right] \text{kJ/mol} \]
\[ - \left[ 1(-174) + 8(0) \right] \text{kJ/mol} \]
\[ = -3512 \text{kJ/mol} \]

and
\[ \Delta S = \left[ 5(214) + 6(70) \right] \text{J/mol} \cdot \text{K} \]
\[ - \left[ 1(263) + 8(205) \right] \text{J/mol} \cdot \text{K} \]
\[ = -413 \text{J/mol} \cdot \text{K} \]
\[ = -0.413 \text{kJ/mol} \cdot \text{K} \]

\[ \Delta G = \Delta H - T \Delta S \]
\[ = -3512 \text{kJ/mol} \]
\[ - (298 \text{K})(-0.413 \text{kJ/mol} \cdot \text{K}) \]
\[ = -3389 \text{kJ/mol} \]

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025 10.0 points

Consider the data below regarding different allotropes of carbon.

<table>
<thead>
<tr>
<th>carbon allotrope</th>
<th>( \Delta G_f^\circ ) kJ \cdot mol(^{-1} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>C(s, graphite)</td>
<td>0</td>
</tr>
<tr>
<td>C(s, diamond)</td>
<td>2.9</td>
</tr>
<tr>
<td>C(_{60})(s, buckminsterfullerene)</td>
<td>24</td>
</tr>
</tbody>
</table>

Which of the following statements is supported by these data?

1. \( C_{60} \) is thermodynamically more stable than graphite under standard conditions.

2. Graphite could spontaneously form \( C_{60} \) under standard conditions.

3. Formation of graphite from \( C_{60} \) would be exergonic under standard conditions. **Correct**

4. Diamond is the least thermodynamically stable allotrope of carbon under standard conditions.

Explanation: