This print-out should have 8 questions. Multiple-choice questions may continue on the next column or page — find all choices before answering.

LDE Stat Thermo Theory 001

001 10.0 points

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

I) An experimentally determined value of \( W \) will be less than the ideal value.

II) For the same number of molecules at absolute zero, the experimental value of \( W \) for HF will be greater than the experimental value for CO.

III) At absolute zero, the positional entropy of a group of molecules with one orientation will equal zero.

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

Explanation:
Statement I is true because ideal values assume that IMF aren’t present and therefore don’t influence the likelihood of a molecule adopting a particular orientation. Statement II is false because HF has a stronger IMF which thus more dramatically perturbing the value of \( W \). Statement III is true because 1 raised to any power is 1 and the natural log of 1 is 0.

LDE Stat Thermo Energy 001

002 10.0 points

Assuming vibrational modes are excited, how much internal energy would be associated with just the vibrational motion of 10 molecules of \( \text{CO}_2 \)?

1. \( 20kT \) correct
2. \( 30kT \)
3. \( 15kT \)
4. \( 40kT \)
5. \( 45kT \)

Explanation:
The energy of any given mode of movement for any given atom or molecule is equal to \( 1/2kT \). Since carbon dioxide has 9 total modes, 3 of which are translational and 2 of which are rotational (because it is linear), it has 4 vibrational modes. Since there are 10 molecules and 4 modes, each of which are worth \( 1/2kT \), we have a total of \( 20kT \) in internal energy.

LDE Stat Thermo Entropy 001

003 10.0 points

If you had 1 mole of HF molecules, each of which had two possible orientations, what would their total theoretical positional entropy be at absolute zero?

1. \( 9.57 \times 10^{-24} \text{ J} \cdot \text{K}^{-1} \)
2. \( 461 \text{ J} \cdot \text{K}^{-1} \)
3. \( 7.65 \times 10^{-22} \text{ J} \cdot \text{K}^{-1} \)
4. \( 5.76 \text{ J} \cdot \text{K}^{-1} \) correct

Explanation:
\( S = k \ln(2^{N_A}) = R \ln 2 = 5.76 \text{ J} \cdot \text{K}^{-1} \)

LDE Internal Energy 001

004 10.0 points

Which of the following statements concerning internal energy is/are true?

I) If the expansion work is small, \( \Delta H \) and \( \Delta U \) are close in value.

II) The internal energy of a system is equal
to $q$ at constant volume.

III) Assuming no heat is exchanged, when pressure-volume work is done on the system, $\Delta U$ is positive.

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

Explanation:
Statement I follows from the identity $\Delta H = \Delta U + p\Delta V$, because $p\Delta V$ is the expansion work. Statement II follows from $\Delta U = q + w$, because $w = 0$ for processes that occur at constant volume. Statement III also follows from $\Delta U = q + w$, because $w$ is the pressure-volume work. (Note: for reversible processes, expansion work and pressure-volume work are identical.)

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**LDE Internal Energy Calc 001**

005 10.0 points

If a man is jogging and releases 800 kJ of heat and does 200 kJ of work, what is the man’s change in internal energy equal to?

1. $-600 \text{ kJ}$
2. $-1000 \text{ kJ}$ correct
3. $1000 \text{ kJ}$
4. $600 \text{ kJ}$

Explanation:
$\Delta U = q + w = -800 \text{ kJ} + (-200 \text{ kJ}) = -1000 \text{ kJ}$

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**LDE Ranking Entropies 001**

006 10.0 points

Rank the following systems in terms of increasing entropy:

a) 1 mol of pure ice
b) 1 mol of water with 1 mol of salt dissolved in it.
c) 1 mol of pure water

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

Explanation:
Entropy increases as systems go through endothermic phase transitions and when there is more matter or more dispersed matter present.

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**LDE Entropy Change Calc 001**

007 10.0 points

If a given system releases 900 J of heat, what will $\Delta S$ of the surroundings be if this takes place at 27 °C?

1. $33.3 \text{ J} \cdot \text{K}^{-1}$
2. $-3 \text{ J} \cdot \text{K}^{-1}$
3. $3 \text{ J} \cdot \text{K}^{-1}$ correct
4. $-33.3 \text{ J} \cdot \text{K}^{-1}$

Explanation:
Heat released by the system will increase the entropy of the surroundings by an amount equal to $q/T$, in this case 

$$\frac{900 \text{ J}}{300 \text{ K}} = 3 \text{ J} \cdot \text{K}^{-1}.$$
Consider the condensation of water at standard pressure at 101 °C. Even without knowing the exact values in $\Delta S_{vap}$ and $\Delta H_{vap}$, what can you say about the value of $\Delta S_{universe}$ for this process?

1. $\Delta S_{universe}$ would be a large positive number.

2. $\Delta S_{universe}$ would be a large negative number.

3. $\Delta S_{universe}$ would be a small negative number. **correct**

4. $\Delta S_{universe}$ would be a small positive number.

**Explanation:**
The process would be non-spontaneous (water doesn’t condense at 1 atm at 101 °C), so $\Delta S_{universe}$ would be negative. We are near the phase change temperature of water, so it would also be small.