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 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 with 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. 9.57×10^{-24} J $\cdot$ K^{-1}

2. 461 J $\cdot$ K^{-1}

3. 7.65×10^{-22} J $\cdot$ K^{-1}

4. 5.76 J $\cdot$ K^{-1} **correct**

Explanation:
   $S = k \ln(2^{N_A}) = R \ln 2 = 5.76$ J $\cdot$ 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, ΔH and ΔU 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, ΔU 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.)

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 kJ

2. -1000 kJ **correct**

3. 1000 kJ

4. 600 kJ

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

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

2. a < c < b **correct**

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.

LDE Entropy Change Calc 001
007 10.0 points

If a given system releases 900 J of heat, what will ΔS 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
$$\frac{900 \text{ J}}{300 \text{ K}} = 3 \text{ J} \cdot \text{K}^{-1}.$$

LDE Thermo 2nd Law Calc 001
008 10.0 points

Consider the condensation of water at standard pressure at 101 °C. Even without knowing the exact values in ΔS_{vap} and ΔH_{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.