1. Which of the following is not one of the laws of thermodynamics?
    1. The entropy of a perfect crystal approaches zero as its temperature approaches zero.
    2. The internal energy of the universe is a constant
    3. The entropy of the universe is increased by all spontaneous processes.
    4. The universe is a closed system.

While the universe is by definition a closed system and this even pertains to the first law, the statement itself is not one of the laws of thermodynamics.

2. If a compound's formation reaction is ________, $\Delta G°_f$ is _______ and the compound is ______.
    1. exothermic, zero, unstable
    2. non-spontaneous, positive, stable
    3. spontaneous, negative, stable
    4. endothermic, negative, stable
    5. at equilibrium, positive, unstable
    6. spontaneous, zero, stable
    7. endothermic, positive, unstable

A spontaneous reaction will have a negative change in free energy and the products will be more stable than the reactants.

3. Which of the following will result in the greatest increase in entropy of the surroundings (ΔS_{surr})?
    1. ΔH_{sys} = -2.68 kJ, T_{surr} = 70 K
    2. ΔH_{sys} = -2,680 J, T_{surr} = 556 K
    3. ΔH_{sys} = 4.13 kJ, T_{surr} = 415 K
    4. ΔH_{sys} = 4,130 J, T_{surr} = 23 K

$\Delta S_{surr} = -\Delta H_{sys}/T_{surr}$

For the change in the thermal entropy of a system's surroundings the be positive (i.e. an increase) the change in enthalpy of the system needs to be negative and the temperature of the surroundings needs to be as small as possible.

4. The enthalpy of fusion (ΔH_{fus}) of tungsten is 35.23 kJ/mol, and its melting point is 3422 °C. What is the entropy of fusion (ΔS_{fus}) of tungsten?
    1. 130.17 J/(mol·K)
    2. 10.30 J/(mol·K)
    3. 10.30 kJ/(mol·K)
    4. 9.53 J/(mol·K)
    5. 4.76 J/(mol·K)

Phase changes are equilibrium processes, and $\Delta G_{fus} = \Delta H_{fus} - T_{fus}\Delta S_{fus} = 0$. Thus:
$\Delta S_{fus} = \Delta H_{fus} / T_{fus}$ = (35,230 J/mol) / (3695 K)

5. Given the enthalpies of formation and molar entropies ($\Delta H°_f$ and $S°_m$) in the following table, what is the Gibbs free energy change (ΔG) for the manufacture of metallic aluminum from alumina at 1300K?

	$\Delta H°_f$ (298 K)	$S°_m$ (298 K)
Al(l)	10.71 $kJ\cdot mol^{-1}$	39.77 $J\cdot mol^{-1}\cdot K^{-1}$
Al_2O_3(l)	-1675.7 $kJ\cdot mol^{-1}$	50.92 $J\cdot mol^{-1}\cdot K^{-1}$
$C_{graphite}$(s)		5.7 $J\cdot mol^{-1}\cdot K^{-1}$
CO_2(g)	-393.51 $kJ\cdot mol^{-1}$	213.74 $J\cdot mol^{-1}\cdot K^{-1}$

$2\,Al_2O_3(l) + 3\,C_{graphite}(s) \rightarrow 4\,Al(l) + 3\,CO_2(g)$ ΔG = ?

1. -1,327.94 kJ·mol^{-1}
2. 1,327.94 kJ·mol^{-1}
3. 1036.94 kJ·mol^{-1}
4. -1036.94 kJ·mol^{-1}

$\Delta H = (4 \times 10.71) + (3 \times -393.51) - (2 \times -1675.7) = 2{,}213.71$ kJ·mol^{-1}
$\Delta S = (4 \times 39.77) + (3 \times 213.74) - (2 \times 50.92) - (3 \times 5.7) = 681.36$ J·mol^{-1}·K^{-1}
$\Delta G = \Delta H - T\Delta S = 2{,}213.71 - 1300(0.68136) = 1{,}327.94$ kJ·mol^{-1}

6. Your roommate left 1 kg of dry ice out on the counter last night and all of it sublimated. Given that ΔH = 393.5 kJ·mol^{-1} and ΔS = 2.023 kJ·mol^{-1}·K^{-1} for dry ice sublimating and that the temperature in the room was 25 °C, by how much in total has has your roommate increased the entropy of the universe?

1. -1.320 kJ·K^{-1}
2. 0 kJ·K^{-1}
3. 0.703 kJ·K^{-1}
4. 2.023 kJ·K^{-1}
5. 15.97 kJ·K^{-1}
6. 45.96 kJ·K^{-1}
7. -39.99 kJ·K^{-1}

$\Delta S_{surr} = -q/T = -\Delta H/T = -(393.5$ kJ·mol$^{-1})/(298$ K$) = -1.320$ kJ·mol^{-1}·K^{-1} (at constant pressure)
$\Delta S_{univ} = \Delta S_{sys} + \Delta S_{surr} = 2.023$ kJ·mol^{-1}·K^{-1} - 1.320 kJ·mol^{-1}·K^{-1} = 0.703 kJ·mol^{-1}·K^{-1}
1 kg of dry ice = (1000 g / 44.010 g/mol) = 22.72moles
so the total increase is (22.72moles)*(0.703 kJ·mol^{-1}·K^{-1}) = 15.97 kJ·K^{-1}