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**

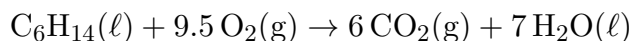**5.** The change in the value of the state function is always positive for endothermic processes.

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

$$C_6H_{14}(\ell) + 9.5\,O_2(g) \rightarrow 6\,CO_2(g) + 7\,H_2O(\ell)$$

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:
$n_i = 9.5$ mol gas $n_f = 6$ mol gas
For $P = $ 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 (ΔU) 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 decreasesed - 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,

$$
\begin{aligned}
w &= -P_{\text{ext}}\, \Delta V \\
&= (-1.2 \text{ atm})(18 \text{ L} - 6 \text{ L}) \\
&\quad \times (101.325 \text{ J} \cdot \text{L}^{-1} \cdot \text{atm}^{-1}) \\
&= -1.45908 \text{ kJ}.
\end{aligned}
$$

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 re-leases 50 kJ as heat due to friction from spin-ning. 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 kJ.

Work *from both together* on the *surroundings* is -250 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 J/g·°C, and of steam 2.03 J/g·°C. The heat of vaporization of water (ℓ) is 2.26 kJ/g and its boiling point is 100°C. What is the total heat flow when 18 grams of water at 12°C are heated to become steam at 109°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:

$$\underset{12°C}{\overset{18\text{ g}}{H_2O(\ell)}} \xrightarrow{\text{step 1}} \underset{100°C}{\overset{18\text{ g}}{H_2O(\ell)}} \xrightarrow{\text{step 2}}$$

$$\underset{100°C}{\overset{18\text{ g}}{H_2O(g)}} \xrightarrow{\text{step 3}} \underset{109°C}{\overset{18\text{ g}}{H_2O(g)}}$$

Step 1: $\dfrac{4.184 \text{ J}}{\text{g} \cdot° \text{ C}} \cdot (18 \text{ g}) \cdot (100 - 12)°\text{C}$
$$= 6{,}627 \text{ J}$$

Step 2: $\dfrac{2.26 \text{ kJ}}{\text{g}} \cdot (18 \text{ g}) \cdot \dfrac{1000 \text{ J}}{1 \text{ kJ}}$
$$= 40{,}680 \text{ J}$$

Step 3: $\dfrac{2.03 \text{ J}}{\text{g} \cdot °\text{C}} \cdot (18 \text{ g}) \cdot (109 - 100)°\text{C}$
$$= 329 \text{ J}$$

Total $=$ 6627 J $+$ 329 J $+$ 40,680 J
$$= 47{,}636 \text{ J} = 47.636 \text{ kJ}$$

009　10.0 points

Refer to the potential energy diagram shown below.
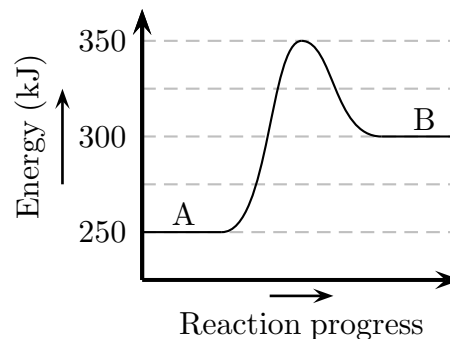
What is the change in enthalpy (ΔH) for the reaction
$$A \rightarrow B \,?$$
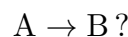
1. $+50$ kJ, endothermic **correct**

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:
$\Delta H_i = \Delta H_A = 250$ kJ
$\Delta H_f = \Delta H_B = 300$ kJ

$$\Delta H = \Delta H_f - \Delta H_i$$
$$= 300 \text{ kJ} - 250 \text{ kJ}$$
$$= 50 \text{ kJ}$$

Notice that ΔH is positive and therefore the reaction is endothermic.

010　10.0 points

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 ΔT 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.

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 ΔU 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. -1.15×10^4 kJ/mol

2. -4.52×10^3 kJ/mol

3. -7.40×10^4 kJ/mol

4. -5.92×10^3 kJ/mol **correct**

5. -9.96×10^3 kJ/mol

Explanation:
$m_{C_6H_8} = 1.00$ g　　　　$m_{water} = 1502$ g
$SH = 4.184$ J/g·°C　　　　$HC = 4042$ J/°C
$\Delta T = 29.30°C - 22.64°C = 6.66°C$

The increase in the water temperature is $29.30°C - 22.64°C = 6.66°C$. The amount of heat responsible for this increase in temperature for 1502 g of water is

$$q = (6.66°C)\left(4.184\frac{J}{g·°C}\right)(1502 \text{ g})$$
$$= 41854 \text{ J} = 41.85 \text{ kJ}$$

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

$$q = (6.66°C)(4042 \text{ J/°C})$$
$$= 26920 \text{ J} = 26.92 \text{ kJ}$$

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.

012　10.0 points
Which of the reactions below is a formation reaction?

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

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

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

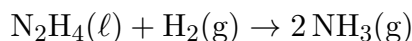**4.** $N_2(\ell) + 2H_2(g) \rightarrow N_2H_4(\ell)$

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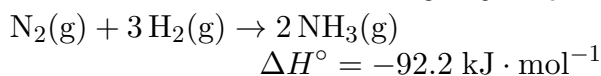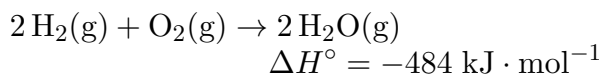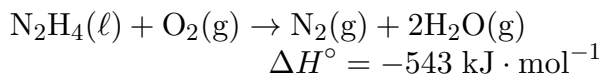

013　10.0 points

Calculate the standard reaction enthalpy for the reaction

$$N_2H_4(\ell) + H_2(g) \rightarrow 2\,NH_3(g)$$

given

$$N_2H_4(\ell) + O_2(g) \rightarrow N_2(g) + 2H_2O(g)$$
$$\Delta H^\circ = -543 \text{ kJ} \cdot \text{mol}^{-1}$$
$$2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g)$$
$$\Delta H^\circ = -484 \text{ kJ} \cdot \text{mol}^{-1}$$
$$N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g)$$
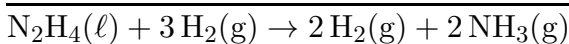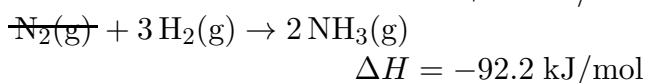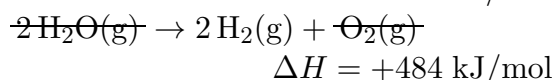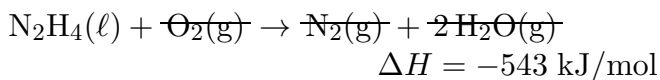$$\Delta H^\circ = -92.2 \text{ kJ} \cdot \text{mol}^{-1}$$

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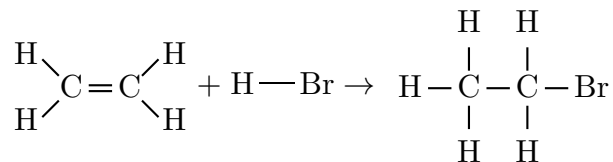We need to reverse the second reaction and add them:

$$N_2H_4(\ell) + \cancel{O_2(g)} \rightarrow \cancel{N_2(g)} + \cancel{2\,H_2O(g)}$$
$$\Delta H = -543 \text{ kJ/mol}$$
$$\cancel{2\,H_2O(g)} \rightarrow 2\,H_2(g) + \cancel{O_2(g)}$$
$$\Delta H = +484 \text{ kJ/mol}$$
$$\cancel{N_2(g)} + 3\,H_2(g) \rightarrow 2\,NH_3(g)$$
$$\Delta H = -92.2 \text{ kJ/mol}$$
$$\overline{N_2H_4(\ell) + 3\,H_2(g) \rightarrow 2\,H_2(g) + 2\,NH_3(g)}$$
$$N_2H_4(\ell) + H_2(g) \rightarrow 2\,NH_3(g)$$
$$\Delta H = -151.2 \text{ kJ/mol}$$

---
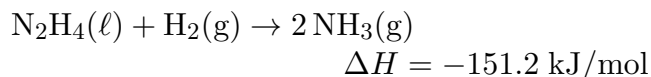
014   10.0 points

Estimate the heat released when ethene $(CH_2\!=\!CH_2)$ reacts with HBr to give CH_3CH_2Br.

1. 200 kJ/mol

2. 76 kJ/mol **correct**

3. 1036 kJ/mol

4. 424 kJ/mol

5. 470 kJ/mol

Explanation:

$$\Delta H = \sum E_{\text{break}} - \sum E_{\text{make}}$$
$$= \Big[(C\!=\!C) + (H\!-\!Br)\Big]$$
$$\quad - \Big[(C\!-\!H) + (C\!-\!Br) + (C\!-\!C)\Big]$$
$$= (602 \text{ kJ/mol} + 366 \text{ kJ/mol})$$
$$\quad - \Big[413 \text{ kJ/mol} + 285 \text{ kJ/mol}$$
$$\quad\quad + 346 \text{ kJ/mol}\Big]$$
$$= -76 \text{ kJ/mol},$$

so 76 kJ/mol of heat was released.

---

015   10.0 points

Consider a reaction that is non-spontaneous at all temperatures. What would be the signs of ΔG_{sys}, ΔH_{surr}, and ΔS_{univ} respectively for such a reaction?

1. +, -, +

2. -, +, -

3. +, -, - **correct**

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.

---

016   10.0 points

Which one shows the substances in the decreasing order of their molar entropy?

1. $C(s)$, $H_2O(\ell)$, $H_2O(g)$, $H_2O(s)$

2. $H_2O(g)$, $H_2O(\ell)$, $H_2O(s)$, $C(s)$ **correct**

3. None of these

4. $H_2O(s)$, $H_2O(\ell)$, $H_2O(g)$, $C(s)$

5. $C(s)$, $H_2O(g)$, $H_2O(\ell)$, $H_2O(s)$

6. $C(s)$, $H_2O(s)$, $H_2O(\ell)$, $H_2O(g)$

Explanation:
Gases will have a higher entropy than liquids so we expect $H_2O(\ell)$ to have the lowest molar entropy. The gases will increase in entropy in the order $Ne(g) < Ar(g) < CO_2(g)$. 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
$H_2O(\ell) < Ne(g) < Ar(g) < CO_2(g)$.

---
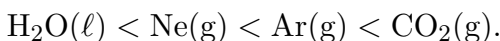
017   10.0 points
Consider the following processes. (Treat all gases as ideal.)
  I) The pressure of one mole of oxygen gas is allowed to double isothermally.
 II) Carbon dioxide is allowed to expand isothermally to 10 times its original volume.
III) The temperature of one mole of helium is increased $25°C$ at constant pressure.
IV) Nitrogen gas is compressed isothermally to one half its original volume.
 V) A glass of water loses 100 J of energy reversibly at $30°C$.
Which of these processes leads to an increase in entropy?

1. III and V

2. I and II

3. I and IV

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 ΔS 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 CO_2 gas volume expanding $10\times$ 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,
$T = 30°C + 273.15 = 303.15 \text{ K}$

$$\Delta S = \frac{q}{T} = \frac{-200 \text{ J}}{303.15 \text{ K}} = -0.6597 \text{ J} \cdot \text{K}^{-1}$$

The last situation (heating the 1 mol of He) does not give enough data to calculate an answer but from the formula

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

$n = 1$ mol and for a monoatomic ideal gas $C_{\text{p, m}} = 2.5\ R$. Finally if the temperature increases this means $T_2 > T_1$ so $\ln\left(\dfrac{T_2}{T_1}\right)$ 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. 1.9×10^{-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_2O) molecules?

1. 15 k T

2. 10 k T

3. 25 k T **correct**

4. 10 R T

5. 25 R T

6. 15 R T

Explanation:
10 molecules times $\dfrac{3}{2}$ k T total translational energy per molecule is $\dfrac{30}{2}$ k T is 15 k T. Since nitrous oxide is a linear molecules, its rotational energy is 10 molecules times k T is 10 k T. The total non-vibrational energy is 25 k T.

---

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.

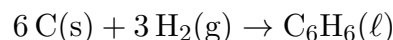**4.** second law of thermodynamics. **correct**

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 <u>any</u> energy conversion 100% efficient.

---

021   10.0 points
Calculate $\Delta S^{\circ}_{\text{surr}}$ at 298 K for the reaction

$$6\,C(s) + 3\,H_2(g) \rightarrow C_6H_6(\ell)$$

$\Delta H^{\circ}_{\text{r}} = +49.0$ kJ·mol^{-1} and $\Delta S^{\circ}_{\text{r}} = -253$ J·K^{-1}·mol^{-1}.

1. -417 J·K^{-1}·mol^{-1}

2. -164 J·K^{-1}·mol^{-1} **correct**

3. $+253$ J·K^{-1}·mol^{-1}

4. $- 253$ J·K^{-1}·mol^{-1}

5. $+164$ J·K^{-1}·mol^{-1}

Explanation:
$\Delta H_r^\circ = 49000$ J $\cdot$ mol^{-1} $T = 298$ K

$$\begin{aligned}
\Delta S_{\text{surr}}^\circ &= \frac{q_{\text{surr}}}{T} = \frac{-q}{T} = \frac{-\Delta H_r^\circ}{298 \text{ K}} \\
&= \frac{-(+49000 \text{ J} \cdot \text{mol}^{-1})}{298 \text{ K}} \\
&= -164.43 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}.
\end{aligned}$$

---

022   10.0 points
For the evaporation of water from an open pan at 25°C, the values of ΔS 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.

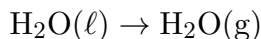**5.** positive, negative, positive. **correct**

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.

$$S(g) > S(\ell) > S(s).$$
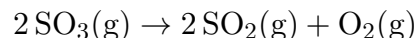
The system is

$$H_2O(\ell) \rightarrow H_2O(g)$$

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

$$2\,SO_3(g) \rightarrow 2\,SO_2(g) + O_2(g)$$

$\Delta H_r^\circ = +198$ kJ·mol^{-1} at 298 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. ΔG_r° will be positive at high temperatures.

5. ΔG_r° will be negative at high temperatures. **correct**

Explanation:
$\Delta G = \Delta H - T\Delta S$ is used to predict spontaneity. (ΔG is negative for a spontaneous reaction.) ΔH 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 ΔS is positive and $\Delta G = (+) - T(+)$. For ΔG to be negative, T must be large.

---

024   10.0 points

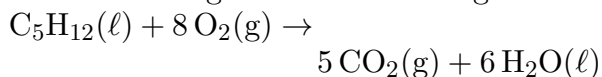What is ΔG_r° for the combustion of liquid n-pentane?

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.
$$C_5H_{12}(\ell) + 8\,O_2(g) \rightarrow$$
$$5\,CO_2(g) + 6\,H_2O(\ell)$$

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

and

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

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

---

025   10.0 points

Consider the date below regarding different allotropes of carbon.

carbon allotrope	$\Delta G_{\text{f}}^{\circ}$ $\text{kJ} \cdot \text{mol}^{-1}$
$C(s, \text{graphite})$	0
$C(s, \text{diamond})$	2.9
$C_{60}(s, \text{buckminsterfullerene})$	24

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: