This print-out should have 21 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

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001 10.0 points

Calculate the change in enthalpy for the combustion of graphite using the data below.

$$H_2(g) + \frac{1}{2}O_2(g) \longleftrightarrow H_2O(l)$$
$$\Delta H = -285.83\ \text{kJ} \cdot \text{mol}^{-1}$$

$$CO_2(g) + 2H_2O(l) \longleftrightarrow CH_4(g) + 2O_2(g)$$
$$\Delta H = 882.00\ \text{kJ} \cdot \text{mol}^{-1}$$

$$C_{\text{graphite}}(s) + 2H_2(g) \longleftrightarrow CH_4(g)$$
$$\Delta H = -74.87\ \text{kJ} \cdot \text{mol}^{-1}$$

$$C_{\text{graphite}}(s) + O_2(g) \longleftrightarrow CO_2(g)$$
$$\Delta H = ?$$

1. $-385.21\ \text{kJ} \cdot \text{mol}^{-1}$ **correct**

2. $-515.87\ \text{kJ} \cdot \text{mol}^{-1}$

3. $521.30\ \text{kJ} \cdot \text{mol}^{-1}$

4. $-230.04\ \text{kJ} \cdot \text{mol}^{-1}$

5. $235.47\ \text{kJ} \cdot \text{mol}^{-1}$

Explanation:
In order for the three provided reactions to cancel to result in the unknown reaction (the combustion of graphite), the first reaction needs to be reversed and doubled, the second needs to be reversed, and the third will remain unchanged. Consequently, the overall change in enthalpy for the reaction is: $\Delta H_{\text{rxn}} = -2 \times -285.83 + -1 \times 882.00 + -74.87 = -385.21\ \text{kJ} \cdot \text{mol}^{-1}$

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002 10.0 points

Which of the reactions below will likely have the largest increase in entropy (ΔS_{rxn})?

1. $C_5H_{12}(l) + 8O_2(g) \rightarrow 6H_2O(g) + 5CO_2(g)$ **correct**

2. $N_2H_4(g) + H_2(g) \rightarrow 2NH_3(g)$

3. $Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$

4. $S_3(g) + 9F_2(g) \rightarrow 3SF_6(g)$

5. $2CH_4(g) + 2O_3(g) \rightarrow 4H_2O(g) + 2CO(g)$

Explanation:
The reaction with the greatest positive value for Δn_{gas} will have the greatest value of ΔS_{rxn}.

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003 10.0 points

What is true about the first law of thermodynamics?

1. $\Delta E_{\text{univ}} < 0$

2. $\Delta E_{\text{univ}} = 0$ **correct**

3. $\Delta E_{\text{sys}} = 0$

4. $\Delta E_{\text{sys}} < 0$

5. $\Delta E_{\text{sys}} > 0$

6. $\Delta E_{\text{univ}} > 0$

Explanation:

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004 10.0 points

Which of

$O_2(g)$, $O_2(\ell)$, $H_2(g)$, $H_2(\ell)$, $H_2O(g)$, $H_2O(\ell)$

have a heat of formation equal to zero?

1. $O_2(g)$, $O_2(\ell)$, $H_2(g)$, $H_2(\ell)$, $H_2O(g)$, $H_2O(\ell)$

2. $O_2(g)$, $O_2(\ell)$, $H_2(g)$, $H_2(\ell)$

3. All of them, but only at absolute zero

4. $O_2(g)$, $H_2(g)$ **correct**

5. $O_2(g)$, $H_2(g)$, $H_2O(g)$

Explanation:

Molecules in their native state at STP have a heat of formation of zero.

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005 10.0 points

If 25.0 g of water at 100.0 °C are mixed with 15.0 g of water at 40.0 °C, what temperature will the 40.0 g of combined water be at once they reach equilibrium?

1. 70.0 °C

2. 77.5 °C **correct**

3. 60.0 °C

4. 62.5 °C

Explanation:

$q_{cold} = -q_{hot}$

$15 \cdot 4.184 \cdot (T_f - 40) = -25 \cdot 4.184 \cdot (T_f - 100)$

$15 \cdot (T_f - 40) = -25 \cdot (T_f - 100)$

$40 \cdot T_f = 3100$

$T_f = 77.5$ °C

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006 10.0 points

Calculate the reaction enthalpy for the formation

$$2\,\mathrm{Al(s)} + 3\,\mathrm{Cl_2(g)} \longrightarrow 2\,\mathrm{AlCl_3(s)}\,,$$

of anhydrous aluminum chloride using the data

$2\,\mathrm{Al(s)} + 6\,\mathrm{HCl(aq)} \longrightarrow 2\,\mathrm{AlCl_3(aq)} + 3\,\mathrm{H_2(g)}$ $\Delta H^\circ = -1049$ kJ

$\mathrm{HCl(g)} \longrightarrow \mathrm{HCl(aq)}$ $\Delta H^\circ = -74.8$ kJ

$\mathrm{H_2(g)} + \mathrm{Cl_2(g)} \longrightarrow 2\,\mathrm{HCl(g)}$ $\Delta H^\circ = -185$ kJ

$\mathrm{AlCl_3(s)} \longrightarrow \mathrm{AlCl_3(aq)}$ $\Delta H^\circ = -323$ kJ

1. −1883.5 kJ

2. −1100.36 kJ

3. −1225.7 kJ

4. −1502.4 kJ

5. −1450.85 kJ

6. −1406.8 kJ **correct**

7. −1826.2 kJ

Explanation:

Using Hess' Law:

The desired reaction is obtained by adding reaction 1; 6 times reaction 2; 3 times reaction 3; and 2 times the reverse of reaction 4:

$$2\,\mathrm{Al(s)} + 6\,\mathrm{HCl(aq)} \longrightarrow 2\,\mathrm{AlCl_3(aq)} + 3\,\mathrm{H_2(g)} \qquad \Delta H^\circ = -1049\ \mathrm{kJ}$$

$$6\,\mathrm{HCl(g)} \longrightarrow 6\,\mathrm{HCl(aq)} \qquad \Delta H^\circ = 6\,(-74.8\ \mathrm{kJ}) = -448.8\ \mathrm{kJ}$$

$$3\,\mathrm{H_2(g)} + 3\,\mathrm{Cl_2(g)} \longrightarrow 6\,\mathrm{HCl(g)} \qquad \Delta H^\circ = 3\,(-185\ \mathrm{kJ}) = -555\ \mathrm{kJ}$$

$$2\,\mathrm{AlCl_3(aq)} \longrightarrow 2\,\mathrm{AlCl_3(s)} \qquad \Delta H^\circ = 2\,(323\ \mathrm{kJ}) = 646\ \mathrm{kJ}$$

$$2\,\mathrm{Al(s)} + 3\,\mathrm{Cl_2(g)} \longrightarrow 2\,\mathrm{AlCl_3(s)} \qquad \Delta H^\circ = -1406.8\ \mathrm{kJ}$$

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007 10.0 points

The pressure-volume work done by an ideal gaseous system at constant volume is

1. $-\dfrac{\Delta P}{P}$

2. zero **correct**

3. $-\Delta E$

4. q

5. $-V\,\Delta P$

Explanation:

When V = constant, nothing moves through a distance and therefore no work is done: $w = 0$.

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008 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. −700 kJ

2. −200 kJ

3. −800 kJ **correct**

4. −750 kJ

5. +200 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.

$$\begin{aligned}\Delta U &= q + w \\ &= [-50\text{ kJ} + (-250\text{ kJ})] + (-500\text{ kJ}) \\ &= -800\text{ kJ}\end{aligned}$$

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009 10.0 points

Which of the following would probably have a positive ΔS value?

1. $\text{He(g, 2 atm)} \rightarrow \text{He(g, 10 atm)}$

2. $H_2(g) + I_2(s) \rightarrow 2\,HI(g)$ **correct**

3. $2\,NO_2(g) \rightarrow N_2O_4(g)$

4. $O_2(g) \rightarrow O_2(aq)$

5. $2\,Ag(s) + Br_2(\ell) \rightarrow 2\,AgBr(s)$

Explanation:

We can predict the sign of ΔS by noting the relative order of entropy: solids (lowest) < liquids < solutions < gases (highest) and the number of moles of each type. Also, if a gas is compressed (pressure increased) its entropy decreases.

For the reactions given we have

$\text{He(g, 2 atm)} \rightarrow \text{He(g, 10 atm)}$

gas pressure increased $\Delta S < 0$

$H_2(g) + I_2(s) \rightarrow 2\,HI(g)$

1 mol gas + 1 mol solid → 2 mol gas; $\Delta S > 0$

$2\,Ag(s) + Br_2(\ell) \rightarrow 2\,AgBr(s)$

2 mol solid + 1 mol liquid → 2 mol solid; $\Delta S < 0$

$O_2(g) \rightarrow O_2(aq)$

1 mol gas → 1 mol solution; $\Delta S < 0$

$2\,NO_2(g) \rightarrow N_2O_4(g)$

2 mol gas → 1 mol gas; $\Delta S < 0$

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010 10.0 points

Which of the following statements is always true?

1. An exothermic reaction is spontaneous.

2. If the number of moles of gas does not change in a chemical reaction, then $\Delta S^0 = 0$.

3. A reaction for which delta S^0 is positive is spontaneous.

4. If ΔH^0 and ΔS^0 are both positive, ΔG^0 will decrease as the temperature increases. **correct**

Explanation:

$$\Delta H = (+)$$
$$\Delta S = (+)$$

$$\begin{aligned}\Delta G &= \Delta H - T\,\Delta S \\ &= (+) - T\,(+) \\ &= (+) - T\end{aligned}$$

As T increases, ΔG decreases.

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011 10.0 points

For the reaction

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

$\Delta H^\circ_r = +198\ \text{kJ}\cdot\text{mol}^{-1}$ and $\Delta S^\circ_r = 190\ \text{J}\cdot\text{K}^{-1}\cdot\text{mol}^{-1}$ at 298 K. The forward reaction will be spontaneous at

1. temperatures above 1042 K. **correct**

2. all temperatures.

3. temperatures above 1315 K.

4. temperatures below 1042 K.

5. no temperature.

Explanation:

$\Delta H_{\mathrm{r}}^{\circ} = +198$ kJ/mol $\quad \Delta S_{\mathrm{r}}^{\circ} = 0.019\ \dfrac{\mathrm{kJ}}{\mathrm{mol}\cdot\mathrm{K}}$

$$\Delta G^0 = \Delta H^0 - T\Delta S$$

$\Delta G^0 < 0$ for a spontaneous reaction, so

$$\begin{aligned} 0 &> \Delta H^0 - T\,\Delta S \\ T &> \frac{\Delta H^0}{\Delta S^0} \\ &= \frac{198\ \mathrm{kJ/mol}}{0.019\ \frac{\mathrm{kJ}}{\mathrm{mol\cdot K}}} = 1042.11\ \mathrm{K} \end{aligned}$$

Thus the temperature would need to be > 1042.11 K.

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012 10.0 points

Which of the following is **not** a definition of enthalpy or change in enthalpy?

1. A measure of the motional energy of a system. **correct**

2. A correction for internal energy that accounts for pressure-volume work.

3. A measure of the heat of a system at constant pressure.

4. A measure of a system's ability to change the entropy of its surroundings.

Explanation:

Total motional energy is a a measure of the internal energy (U, or sometimes E) of a system, not the enthalpy (H).

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013 10.0 points

Which of the following quantities is **not** path independent?

1. entropy (S)

2. volume (V)

3. pressure (P)

4. heat (q) **correct**

5. enthalpy (H)

Explanation:

Heat quantitatively describes a transition **between** states, but not a state itself, and is thus a process quantity (path function), not a state function.

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014 10.0 points

Which of the following would experience the smallest increase in temperature if 1 kJ of heat were added to it?

1. 1 g of copper metal

2. 10 g of water **correct**

3. 10 g of copper metal

4. 1 g of water

Explanation:

Heat capacity is an extensive property, that reflects the amount of energy required to raise the temperature of an object. Because water has a much higher heat capacity than metals (to some extent a consequence of IMF), and because 10 grams is more than 1 gram, the 10 gram sample of water would experience a much smaller increase in temperature than the other choices.

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015 10.0 points

Given the following data:

$3CO_2(g) + 4H_2O(l) \longleftrightarrow C_3H_8(g) + 5O_2(g)$
$\Delta H = 1,110\ \mathrm{kJ\cdot mol^{-1}}$

$H_2O(l) \longleftrightarrow H_2(g) + 1/2O_2(g)$
$\Delta H = 142.5\ \mathrm{kJ\cdot mol^{-1}}$

$3C_{\mathrm{graphite}} + 4H_2(g) \longleftrightarrow C_3H_8(g)$

$\Delta H = -52\ \text{kJ} \cdot \text{mol}^{-1}$

calculate ΔH for the reaction

$C_{\text{graphite}} + O_2(g) \longleftrightarrow CO_2(g)$

1. $-592\ \text{kJ} \cdot \text{mol}^{-1}$

2. $-543\ \text{kJ} \cdot \text{mol}^{-1}$

3. $-197\ \text{kJ} \cdot \text{mol}^{-1}$ **correct**

4. $543\ \text{kJ} \cdot \text{mol}^{-1}$

5. $1202\ \text{kJ} \cdot \text{mol}^{-1}$

Explanation:

$$\Delta H_{rxn} = \frac{11}{3}(1,110\ \text{kJ} \cdot \text{mol}^{-1}) - \frac{4}{3}(142.5\ \text{kJ} \cdot \text{mol}^{-1}) - \frac{1}{3}(-52\ \text{kJ} \cdot \text{mol}^{-1}) = -197\ \text{kJ} \cdot \text{mol}^{-1}$$

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016 10.0 points

Rank the following reactions from least to greatest in terms of change in entropy (ΔS_{rxn}):

a) $KNO_3(aq) + NaCl(aq) \longrightarrow NaNO_3(aq) + KCl(aq)$
b) $4Ag(s) + O_2(g) \longrightarrow 2Ag_2O(s)$
c) $2H_2O_2(\ell) \longrightarrow 2H_2O(g) + O_2(g)$
d) $NaHCO_3(s) \longrightarrow NaOH(s) + CO_2(g)$

1. b < a < c < d

2. a < b < d < c

3. c < d < b < a

4. b < a < d < c **correct**

5. a < d < b < c

6. d < c < a < b

Explanation:

Reaction b has a decrease in total moles and a decrease in moles of gas; ΔS_{rxn} is a negative number. In reaction a, no changes in phase and no changes in number of moles occur; ΔS_{rxn} will be nearly zero. Reaction d evolves a single mole of gas; ΔS_{rxn} will be a small positive number. Reaction c evolves 3 moles of gas; ΔS_{rxn} will be a large positive number.

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017 10.0 points

Phosphine (the common name for PH3, a highly toxic gas used for fumigation), has a $\Delta H^\circ_{vap} = 14.6\ \text{kJ} \cdot \text{mol}^{-1}$ and a $S^\circ_{vap} = 78.83\ \text{J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$. What is the normal boiling point of phosphine expressed in centigrade?

1. 185.2 °C

2. −0.2 °C

3. −87.8 °C **correct**

4. 273 °C

Explanation:

Because boiling is an equilibrium process, $\Delta G^\circ_{vap} = 0 = \Delta H^\circ_{vap} - T\Delta S^\circ_{vap}$.

And so $T\Delta S^\circ_{vap} = \Delta H^\circ_{vap}$
and $T = \Delta H^\circ_{vap}/\Delta S^\circ_{vap}$
$= 14,600\text{J} \cdot \text{mol}^{-1}/78.83\ \text{J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$
$= 185.2$ K
$= -87.8$ °C

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018 10.0 points

Which of the following combustion reactions does not occur at high temperature?

I) $C_6H_{12}O_6(s) + 6O_2(g) \longleftrightarrow 6CO_2(g) + 6H_2O(g)$
II) $2H_2(g) + O_2(g) \longleftrightarrow 2H_2O(g)$
III) $C_3H_8(g) + 5O_2(g) \longleftrightarrow 3CO_2(g) + 4H_2O(g)$

1. II, III

2. II only **correct**

3. III only

4. I, II

5. I, II, III

6. I, III

7. I only

Explanation:

Only reaction II results in a decrease in entropy, and is thus becomes less spontaneous and eventually non-spontaneous as temperature increases.

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019 10.0 points

(For this problem assume that the calorimeter itself absorbs no heat, the density of water is $1.00\ \mathrm{g \cdot mL^{-1}}$*, and the specific heat capacity of water is* $4.184\ \mathrm{J \cdot g^{-1} \cdot K^{-1}}$*.)*

1.14 g of octane (C_8H_{18}) is combusted in a bomb calorimeter surrounded by 1 L of water. The initial and final temperatures of the water are 25 °C and 38 °C respectively. Determine the molar enthalpy of combustion of octane.

1. $-5,440\ \mathrm{kJ \cdot mol^{-1}}$ **correct**

2. $-54,400\ \mathrm{kJ \cdot mol^{-1}}$

3. $-54.4\ \mathrm{kJ \cdot mol^{-1}}$

4. $-544\ \mathrm{kJ \cdot mol^{-1}}$

5. $-5.44\ \mathrm{kJ \cdot mol^{-1}}$

Explanation:

$\Delta T = T_f - T_i = 38\ ^\circ\mathrm{C} - 25\ ^\circ\mathrm{C} = 13\ ^\circ\mathrm{C} = 12.98\ \mathrm{K}$

$$m = 1\ \mathrm{L} \cdot \frac{1000\ \mathrm{mL}}{\mathrm{L}} \cdot \frac{1.00\ \mathrm{g}}{\mathrm{mL}} = 1000\ \mathrm{g}$$

$$n = 1.14\ \mathrm{g\ octane} \cdot \frac{114\ \mathrm{g}}{\mathrm{mol}} = 0.01\ \mathrm{mol}$$

$$\Delta H_{rxn} = -\Delta H_{cal} = -mc\Delta T$$

$$\frac{-mc\Delta T}{n} = -\frac{1000\ \mathrm{g} \cdot \frac{4.184\ \mathrm{J}}{\mathrm{g \cdot K}} \cdot 13\ \mathrm{K}}{0.01\ \mathrm{mol}}$$

$$= -5,440\ \mathrm{kJ \cdot mol^{-1}}$$

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020 10.0 points

If both Δn_{gas} and Δn for a particular reaction are zero, which statement below would be the most accurate prediction we could make about the change in entropy?

1. It is definitely positive.

2. It may be either positive or negative and is probably very small. **correct**

3. It is definitely negative.

4. It may be either positive or negative and is probably very large.

Explanation:

Because the best way to predict the sign and magnitude of the change in entropy is based on the change in the moles of reactants and products (particularly of gaseous reactants or products), the sign would be impossible to predict, but the change would likely be very small.

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021 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) ΔU for 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, II, III **correct**

2. II only

3. III only

4. I, III

5. I only

6. II, III

7. I, II

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.)