This print-out should have 8 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering. The due time is Central time.

Msci 15 0108
19:03, general, multiple choice, > 1 min, .
001
If a system absorbs heat and also does work on its surroundings, its energy

1. must increase.

2. must decrease.

3. must not change.

4. may either increase or decrease, depending on the relative amounts of heat absorbed and work done. **correct**

Explanation:

$$\Delta E = q + w$$

$q > 0$ because heat is absorbed and $w < 0$ because the system does work on its surroundings. Therefore $\Delta E = (+) + (-)$. ΔE can be positive only if $q > w$, and negative only if $w > q$.

ChemPrin3e T06 14
19:03, general, multiple choice, < 1 min, .
002
A system had 150 kJ of work done on it and its internal energy increased by 60 kJ. How much energy did the system gain or lose as heat?

1. The system lost 90 kJ of energy as heat. **correct**

2. The system lost 210 kJ of energy as heat.

3. The system gained 60 kJ of energy as heat.

4. The system gained 90 kJ of energy as heat.

5. The system gained 210 kJ of energy as heat.

Explanation:

DAL Thermo Instability
20:06, general, multiple choice, > 1 min, .
003
Consider the following compounds and their thermodynamic data:

Compound	ΔH_f° $\left(\frac{\text{kJ}}{\text{mol}}\right)$	S° $\left(\frac{\text{J}}{\text{mol}\cdot\text{K}}\right)$	ΔG_f° $\left(\frac{\text{kJ}}{\text{mol}}\right)$
CH_4	-75	186	-50
CH_2O	-108	218	-102
$C_6H_5NH_2$	87	166	319
C_2H_4	52	68	219

Using this data, which of the following answers includes the compounds that are thermodynamically unstable?

1. CH_4, CH_2O, C_2H_4

2. CH_2O, $C_6H_5NH_2$

3. CH_4, C_2H_4

4. $C_6H_5NH_2$, C_2H_4 **correct**

5. Cannot be determined from the data provided.

6. All of the compounds are thermodynamically stable.

Explanation:

ChemPrin3e T07 15
20:04, general, multiple choice, < 1 min, .
004
The enthalpy of fusion of $H_2O(s)$ at its normal melting point is 6.01 kJ $\cdot$ mol^{-1}. What is the entropy change for freezing 1 mole of water at this temperature?

1. $+20.2$ J $\cdot$ K$^{-1}\cdot$ mol^{-1}

2. 0 J $\cdot$ K$^{-1}\cdot$ mol^{-1}

3. $-$ 20.2 J $\cdot$ K$^{-1}\cdot$ mol^{-1}

4. $+22.0\ \mathrm{J \cdot K^{-1} \cdot mol^{-1}}$

5. $-\ 22.0\ \mathrm{J \cdot K^{-1} \cdot mol^{-1}}$ **correct**

Explanation:

ChemPrin3e T07 04a
20:04, general, multiple choice, < 1 min, .
005

The temperature of 2.00 mol Ne(g) is increased from 25°C to 200°C at constant pressure. Assume the heat capacity of Ne is 20.8 J/K-mol. Calculate the change in the entropy of neon. Assume ideal behavior.

1. $+7.68\ \mathrm{J{\cdot}K^{-1}}$

2. $+19.2\ \mathrm{J{\cdot}K^{-1}}$ **correct**

3. $-\ 7.68\ \mathrm{J{\cdot}K^{-1}}$

4. $-\ 19.2\ \mathrm{J{\cdot}K^{-1}}$

5. $+9.60\ \mathrm{J{\cdot}K^{-1}}$

Explanation:

Msci 15 1509
20:06, general, multiple choice, > 1 min, .
006

If you have an endothermic process in which the change in entropy is positive, how can you make it spontaneous?

1. Increasing the pressure

2. Decreasing the volume

3. Increasing the temperature **correct**

4. Decreasing the temperature

5. Reducing the entropy change

Explanation:

$$\Delta G = \Delta H - T\,\Delta S$$

$\Delta H > 0$ for endothermic processes.
$\Delta G < 0$ for spontaneous processes.

T is always positive, so

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

ΔG is negative if T is very large, so increasing the temperature makes the process endothermic.

Msci 15 1412
20:05, general, multiple choice, > 1 min, .
007

Calculate ΔG at 298 K for the reaction

$$2\,\mathrm{Ag_2O(s)} \rightarrow 4\,\mathrm{Ag(s)} + \mathrm{O_2(g)}\,.$$

Species	ΔH_f^0 kJ/mol	S^0 J/mol·K
Ag(s)	0.0	42.55
$Ag_2O(s)$	−30.57	121.7
$O_2(g)$	0.0	205.0

1. 21.9 kJ/mol rxn **correct**

2. 38.2 kJ/mol rxn

3. 52.7 kJ/mol rxn

4. −69.85 kJ/mol rxn

5. 81.2 kJ/mol rxn

Explanation:

$$\begin{aligned}\Delta H^0_\mathrm{rxn} &= \sum n\,\Delta H^0_{\mathrm{f\,prod}} - \sum n\,\Delta H^0_{\mathrm{f\,rct}}\\ &= 0\ \mathrm{kJ/mol}\\ &\quad - 2(-30.57\ \mathrm{kJ/mol})\\ &= 61.14\ \mathrm{kJ/mol}\end{aligned}$$

$$\begin{aligned}\Delta S^0_\mathrm{rxn} &= \sum n\,\Delta S^0_{\mathrm{f\,prod}} - \sum n\,\Delta S^0_{\mathrm{f\,rct}}\\ &= \Big[4(42.55\ \mathrm{J/mol \cdot K})\\ &\qquad + (205.0\ \mathrm{J/mol \cdot K})\Big]\end{aligned}$$

$$
\begin{aligned}
&- 2(121.7\ \mathrm{J/mol\cdot K})\\
&= 131.8\ \mathrm{J/mol\cdot K}\cdot\frac{\mathrm{kJ}}{1000\ \mathrm{J}}\\
&= 0.1318\ \mathrm{kJ/mol\cdot K}
\end{aligned}
$$

$$
\begin{aligned}
\Delta G &= \Delta H - T\,\Delta S\\
&= (+61.14\ \mathrm{kJ/mol})\\
&\quad - (298\ \mathrm{K})(0.1318\ \mathrm{kJ/mol\cdot K})\\
&= 21.8636\ \mathrm{kJ/mol\ rxn}
\end{aligned}
$$

ChemPrin3e T07 42

20:04, general, multiple choice, < 1 min, .

008

The entropy of fusion of water is +22.0 $\mathrm{J\cdot K^{-1}\cdot mol^{-1}}$ and the enthalpy of fusion of water is +6.01 $\mathrm{kJ\cdot mol^{-1}}$ at 0°C. What is $\Delta S_{\mathrm{total}}$ for the melting of ice at 0°C?

1. $-6010\ \mathrm{J\cdot K^{-1}\cdot mol^{-1}}$

2. 0 **correct**

3. $-22.0\ \mathrm{J\cdot K^{-1}\cdot mol^{-1}}$

4. $+6010\ \mathrm{J\cdot K^{-1}\cdot mol^{-1}}$

5. $+22.0\ \mathrm{J\cdot K^{-1}\cdot mol^{-1}}$

Explanation: