001 version ____________ last name ____________ first name ____________________ signature

McCord CH301
unique: 49885
TTh 9:30 am - 11:00 am

Exam 4

Dec 3, 2018
Monday 7:30 - 9:00 PM
A - Mi in BUR 106
Mo - Z in JES A121A

Remember to refer to the Periodic Table handout that is separate from this exam copy.

NOTE: Please keep this exam copy intact (all pages still stapled - including this cover page). You must turn in ALL the materials that were distributed. This means that you turn in your exam copy (name and signature included), bubble sheet, periodic table handout, and all scratch paper. Please also have your UT ID card ready to show as well.

This print-out should have 20 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

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001 5.0 points

Which one of the following thermodynamic quantities is NOT a state function?

1. pressure

2. entropy

3. temperature

4. free energy

5. heat **correct**

Explanation:

All except heat are independent of path and are dependent solely on the current state of the system. Heat (as well as work) is not a state function. It is dependent on the process or pathway. HOW you got there DOES affect the value of heat and work.

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002 5.0 points

WITHDRAWN

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003 5.0 points

Calculate the value of work when a gas-phase reaction compresses from 43.7 L to 36.3 L at 1.85 atm and constant temperature. Is work done on or by the system?

1. 1387 J, on the system **correct**

2. -1387 J, on the system

3. 13.69 J, on the system

4. 0.1351 J, by the system

5. -0.1351 J, on the system

6. 1387 J, by the system

Explanation:

$$\begin{aligned}\text{work} &= -P\Delta V = -(1.85\,\text{atm})(43.7\,\text{L} - 36.3\,\text{L})\\ &= 13.69\,\text{L}\cdot\text{atm}\end{aligned}$$

$$\text{work} = 13.69\,\text{L}\cdot\text{atm} \times \frac{101.325\,\text{J}}{1\,\text{L}\cdot\text{atm}} = 1387\,\text{J}$$

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004 5.0 points

A 14.0 g sample of ice at −18.0°C is placed on a hot plate and heated to a final temperature of 84.0°C. Calculate the heat of this process.

1. 16.2 kJ

2. -11.0 kJ

3. 10.7 kJ

4. 10.1 kJ **correct**

5. 5.8 kJ

Explanation:

You will need to make three different calculations for this problem:

1) Heat ice from -18°C to 0°C:

$$\begin{aligned}q = mC\Delta T &= (14\,\text{g})(2.09\,\frac{\text{J}}{\text{g}^\circ\text{C}})[0^\circ\text{C} - (-18\,^\circ\text{C})]\\ &= 526.68\,\text{J}\end{aligned}$$

2) Melt the ice at constant temperature:

$$\begin{aligned}q = (m)(\Delta H_{\text{fus}}) &= (14\,\text{g})(334\,\frac{\text{J}}{\text{g}})\\ &= 4676\,\text{J}\end{aligned}$$

3) Heat the water from 0°C to 84°C:

$$\begin{aligned}q = mC\Delta T &= (14\,\text{g})(4.184\,\frac{\text{J}}{\text{g}^\circ\text{C}})[-18^\circ\text{C} - 0^\circ\text{C}]\\ &= 4920.384\,\text{J}\end{aligned}$$

Total heat required:

$(526.68\,\text{J}) + (4676\,\text{J}) + (4920.384\,\text{J}) = 10123.064\,\text{J}$

Or:

$= 10.1$ kJ

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005 5.0 points

The atmospheric photochemical oxidation of nitrogen dioxide is an important contributor to rising HNO_3 levels in coastal acid rains. To simulate the first steps of this process, you run the following reaction in a piston:

$$2NO_2(g) + NaCl(s) \rightarrow NOCl(g) + NaNO_3(s)$$

Which of the following is true regarding this reaction?

1. $\Delta U_{\text{sys}} < \Delta H_{\text{sys}}$

2. $\Delta U_{\text{sys}} = \Delta nRT$

3. Work is done by the system

4. $\Delta U_{\text{sys}} = \Delta H_{\text{sys}}$

5. $\Delta U_{\text{sys}} > \Delta H_{\text{sys}}$ **correct**

Explanation:
Based on the first law equation:

$$\Delta\, U = \Delta\, H - \Delta\, nRT,$$

It can be seen that the work term is positive (n = -1, work is done on the system via compression). Mathematically, this means that $\Delta U_{\text{sys}} > \Delta H_{\text{sys}}$.

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006 5.0 points

When a 0.6890 g sample of a petroleum extract (molecular weight = 62.15 g/mol) is combusted in a rigid container, the temperature increases from 25.317°C to 28.918°C. The total volume of water is 1.018 L. The sum of all hardware components of the calorimeter have a heat capacity of 1.78 kJ/°C. Calculate the internal energy of combustion for one mole of this petroleum extract.

1. +4.416 kJ/mol

2. -1962 kJ/mol **correct**

3. +21.75 kJ/mol

4. +31.56 kJ/mol

5. -4.416 kJ/mol

6. +6420 kJ/mol

7. +1962 kJ/mol

8. -21.75 kJ/mol

Explanation:
Heat released by the combustion reaction:

$$\text{q} = \text{mC}\Delta\,\text{T} + \text{C}\Delta\,\text{T}$$

$$\text{q}_{\text{cal}} = (1018)(4.184)(3.601) + (1.78 \times 1000)(3.601)$$

$$\text{q}_{\text{cal}} = 21747.6\,\text{J}$$

For the system, q_v = -q_{cal}:

$$\text{q}_\text{v} = -21.7476\,\text{kJ}$$

This is how much heat was released when 0.689 g of hydrocarbon was combusted (we consider it to be negative). Convert to this to how much heat would be released per mole of ethylene combusted.

$\frac{-21.7476\,\text{kJ}}{0.689\,\text{g}} \times \frac{62.15\,\text{g}}{1\,\text{mol}} = -1962\,\text{kJ/mol}$

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007 5.0 points

The enthalpy change, ΔH, associated with the following reaction is +81 kJ/ mol rxn.

$$NBr_3(g) + 3\,H_2O(g) \rightleftharpoons 3\,HOBr(g) + NH_3(g)$$

What is the expected enthalpy change for the reverse reaction of nine moles of HOBr and two moles of NH_3?

1. -162 kJ **correct**

2. -81 kJ

3. +243 kJ

4. -243 kJ

5. +81 kJ

6. +365 kJ

7. -365 kJ

8. +162 kJ

Explanation:

The reverse reaction is happening when HOBr and NH_3 are mixed. However, they are not mixed in stoichiometric amounts. The 2 moles of NH_3 would require only 6 moles of HOBr. Therefore the HOBr is in excess and the NH_3 is the limiting reactant.

Now you only need to double the ΔH for the reaction and change the sign to negative. This gives you -162 kJ.

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008 5.0 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. $-59\ \text{kJ}\cdot\text{mol}^{-1}$

2. $-151\ \text{kJ}\cdot\text{mol}^{-1}$ **correct**

3. $-1119\ \text{kJ}\cdot\text{mol}^{-1}$

4. $-243\ \text{kJ}\cdot\text{mol}^{-1}$

5. $-935\ \text{kJ}\cdot\text{mol}^{-1}$

Explanation:

We need to reverse the second reaction and add them:

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

$$N_2H_4(\ell) + 3\,H_2(g) \rightarrow 2\,H_2(g) + 2\,NH_3(g)$$
$$N_2H_4(\ell) + H_2(g) \rightarrow 2\,NH_3(g)$$
$$\Delta H = -151.2\ \text{kJ/mol}$$

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009 5.0 points

Chlorine monofluoride (ClF) will react with carbon monoxide (CO) to give carbonyl chlorofluoride (COClF):

$$\text{ClF} + \text{CO} \longrightarrow \text{COClF}$$

Use bond energies (provided elsewhere) to estimate the change in enthalpy (ΔH) for this reaction.

1. -193 kJ/mol

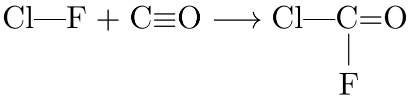
2. -376 kJ/mol

3. -571 kJ/mol

4. -444 kJ/mol

5. -298 kJ/mol **correct**

Explanation:



$$\text{Cl—F} + \text{C}\equiv\text{O} \longrightarrow \text{Cl—}\underset{\underset{\text{F}}{|}}{\text{C}}\text{=O}$$

Break: Cl—F and C≡O
253 1072

Make: C=O and C—Cl and C—F
799 339 485

breaking : +1325 (reactants)
making : -1623 (products)

$\Delta H = -298$ kJ/mol

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010 5.0 points

Which of the following is a formation reaction where $\Delta H^\circ_{\text{rxn}} = \Delta H^\circ_{\text{f}}$?

1. $SO_2(g) + \frac{1}{2}\,O_2(g) \rightarrow SO_3(g)$

2. $Na(s) + \frac{1}{2}\,Cl_2(g) \rightarrow NaCl(s)$ **correct**

3. $CO_2(s) \rightarrow CO_2(g)$

4. $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(\ell)$

Explanation:

All of the reactants in a formation reaction must be ELEMENTS in their thermochemical standard state. AND, to match table values, only 1 MOLE of product can be formed because standard table values are in kJ/mol. Only the NaCl fits this definition. The H_2O reaction is forming 2 moles of water instead of one.

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011 5.0 points

Mercury has a melting point substantially below room temperature (-38.89°C). Which of the following is true regarding the fusion of mercury at room temperature?

I. $\Delta S_{\text{universe}} > 0$

II. $\Delta S_{\text{sys}} > 0$

III. $|\Delta S_{\text{sys}}| > |\Delta S_{\text{surr}}|$

IV. This physical reaction proceeds only at temperatures below -38.89°C

1. I and IV only

2. III and IV only

3. I only

4. II only

5. I, II, and III only **correct**

6. I, II, and IV only

7. IV only

Explanation:

I. True. This reaction is spontaneous. According to the second law of thermodynamics, a spontaneous change has a positive entropy change for the universe.

II. True. Transitioning from a solid to liquid is a positive change in entropy.

III. True. For a physical change, a positive change in entropy for the system is matched by a negative entropy change for the surroundings. The only way to make the entropy change for the universe positive is to ensure that the magnitude of the entropy change of the surroundings is less than entropy change for the system.

IV. False. The reaction only proceeds at temperatures greater than -38.89°C.

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012 5.0 points

Rank the following substances by increasing absolute standard entropy (S°):

$CH_2Cl_2(g)$ $Al(s)$ $CH_3OH(\ell)$ $CH_4(g)$

1. $CH_2Cl_2(g) < CH_4(g) < CH_3OH(\ell) < Al(s)$

2. $CH_4(g) < CH_2Cl_2(g) < CH_3OH(\ell) < Al(s)$

3. $Al(s) < CH_3OH(\ell) < CH_4(g) < CH_2Cl_2(g)$ **correct**

4. $Al(s) < CH_4(g) < CH_2Cl_2(g) < CH_3OH(\ell)$

5. $Al(s) < CH_3OH(\ell) < CH_2Cl_2(g) < CH_4(g)$

Explanation:

The absolute entropy of a substance can be related to the phase and complexity of the molecule. First rank solid, liquid, gas and then determine the relative complexities of CH_4 versus CH_2Cl_2.

This gives the order:

$$Al(s) < CH_3OH(\ell) < CH_4(g) < CH_2Cl_2(g)$$

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013 5.0 points

Calculate the change in entropy (ΔS) when 13 moles of neon gas are heated from 25°C to 125°C at constant pressure.

1. +46.9 J/K

2. +31.26 J/K

3. +27.0 kJ/K

4. +93.8 J/K

5. +3.61 J/K

6. +125 J/K

7. +6.01 J/K

8. +78.2 J/K **correct**

Explanation:

For a monatomic ideal gas:
$C_\text{p} = 20.785$ J/mol K
$\Delta S = nC_\text{p} \ln\left(\frac{T_\text{f}}{T_\text{i}}\right)$
$\Delta S = 13 \cdot 20.785 \ln\left(\frac{398.15}{298.15}\right)$
$\Delta S = 13 \cdot 20.785 \ln(1.335)$
$\Delta S = 13 \cdot 6.012$
$\Delta S = 78.2$ J/K

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014 5.0 points

Calculate the ΔS_surr for the formation of iron(II) oxide at 25°C and 1 atm.

$\text{Fe(s)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{FeO(s)}$

1. 912 J/K **correct**

2. 10.9 J/K

3. -912 J/K

4. -10900 J/K

5. 5440 J/K

6. -10.9 J/K

Explanation:

In general for any process:

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

This is because the heat flow in the surroundings is just the opposite of the heat flow for the system ($q_\text{surr} = -q_\text{sys}$) and at constant pressure the heat is equal to ΔH.

therefore $\Delta S_\text{surr} = 272000/(25 + 273.15)$
$= 912$ J/K

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015 5.0 points

Solid arsenic will sublime when heated to 614°C according to the following reaction:

$\text{As(s)} \rightarrow \text{As(g)}$

What are the signs of ΔH, ΔS, and ΔG for this sublimation reaction at 800°C?

1. $\Delta H > 0, \Delta S > 0, \Delta G = 0$

2. $\Delta H > 0, \Delta S < 0, \Delta G = 0$

3. $\Delta H > 0, \Delta S > 0, \Delta G < 0$ **correct**

4. $\Delta H > 0, \Delta S > 0, \Delta G > 0$

5. $\Delta H < 0, \Delta S < 0, \Delta G < 0$

Explanation:

The phase change from (s) to (g) requires energy, which is a positive change in enthalpy. Creating a gas from a solid is a positive change in entropy. Because this is a positive change in enthalpy and a positive change in entropy, the reaction will be favored ($\Delta G < 0$) at temperatures higher than 614°C.

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016 5.0 points

Consider the equation

$$\text{NH}_4\text{Br(s)} \rightarrow \text{NH}_3\text{(g)} + \text{HBr(g)}$$

carefully, and think about the sign of ΔS for the reaction it describes. $\Delta H = +188.3$ kJ. Which response describes the thermodynamic spontaneity of the reaction?

1. All responses are correct.

2. The reaction is spontaneous at all temperatures.

3. The reaction is spontaneous only at relatively low temperatures.

4. The reaction is spontaneous only at relatively high temperatures. **correct**

5. The reaction is not spontaneous at any temperatures.

Explanation:

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.

$$S(g) > S(\ell) > S(s)\,.$$

A reaction is spontaneous only when ΔG is negative. ΔH is positive for this reaction and ΔS is positive.

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

ΔG will be negative (spontaneous reactions) only at high values of T.

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017 5.0 points

An unidentified metal, M, oxidizes at room temperature and pressure to form M_2O_3. Given the thermodynamic data provided in the table below, calculate $\Delta\, G^\circ_\mathrm{f}$.

Substance	$\Delta H^\circ_\mathrm{f}$ kJ/ mol	S° J/ mol K
M(s)	–	36.2
O_2(g)	–	205
O(g)	249	161
M_2O_3(s)	−653	67.8

1. -560 kJ mol^{-1} **correct**

2. 1250 kJ mol^{-1}

3. 653 kJ mol^{-1}

4. −5430 kJ mol^{-1}

5. 560 kJ mol^{-1}

6. 6.1×10^6 kJ mol^{-1}

7. −653 kJ mol^{-1}

8. −1250 kJ mol^{-1}

Explanation:

For the formation of one mole of M_2O_3, you can use the chemical equation shown below:

$$2M(s) + \frac{3}{2}O_2(g) \rightarrow M_2O_3(s)$$

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

Using $\Delta H^\circ_\mathrm{f}$ and S° values:

$\Delta S = 67.8 - [(1.5 \times 205) + (2 \times 36.2)]$

$$\Delta S = -312.1\,\mathrm{J/K}$$

$$\Delta H = -653\,\mathrm{kJ}$$

$$\Delta G = -560\,\mathrm{kJ}$$

This was solved per mole based on the formation reaction.

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018 5.0 points

Consider a spontaneous, exothermic reaction that has a negative change in entropy. Which of the following relationships is/are true for this reaction?

I. $\Delta S_\mathrm{universe} < 0$

II. $\Delta H_\mathrm{sys} < 0$

III. $\Delta G_\mathrm{rxn} < 0$

1. I, II, and III

2. III only

3. I and III only

4. II only

5. I only

6. II and III only **correct**

Explanation:

I. False. This reaction is spontaneous; therefore the entropy change of the universe must be positive.

II. True. The reaction is exothermic.

III. True. The reaction is spontaneous

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019 5.0 points

Consider the following reaction for the oxidation of lead:

$$Pb(s) + O_2(g) \longrightarrow PbO_2(s)$$

$$\Delta G^\circ_{\text{rxn}} = -217\,\text{kJ/mol}$$

Read each answer choice carefully and determine the single best explanation for the value of $\Delta G^\circ_{\text{rxn}}$.

1. PbO_2 tends to spontaneously decompose into Pb and O_2.

2. PbO_2 has a lower free energy state than that of the Pb and O_2 mixture. **correct**

3. The PbO_2 product has a higher entropy than the combination of Pb and O_2.

4. As the reaction proceeds, work is done by the system to produce a lower free energy.

Explanation:
$\Delta G^\circ_{\text{rxn}}$ indicates that the products have a lower free energy state than the reactants. All other statements are false.

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020 5.0 points

Consider the following combustion reaction:

$$2C_3H_8O(\ell) + 9O_2(g) \rightarrow 6CO_2(g) + 8H_2O(\ell)$$

Which of the following statements about q and w for this reaction is correct?

1. q and w are both are negative.

2. q and w are both zero.

3. q is negative and w is positive. **correct**

4. q is positive and w is negative.

5. q and w are both are positive.

Explanation:
This reaction is a combustion reaction which releases heat (q is negative). The balanced reaction also results in fewer gas moles as products than the reactant moles ($\Delta n < 0$) which means work is positive (work done on the system).