425
version   last name   first name   signature

McCord CH301   # Exam 5   Dec 5, 2017
50070   BUR 106   Tuesday
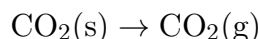TTh 9:30 am - 11 pm

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.

001   5.0 points

Carbon dioxide is a unique substance that sublimes at atmospheric pressure. Assuming no excess heat is supplied to the system at constant pressure, what are the signs of ΔH, ΔS, and w for the following phase change reaction at -78.5°C?

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

1. $\Delta H < 0, \Delta S < 0, w = 0$

2. $\Delta H > 0, \Delta S < 0, w < 0$

3. $\Delta H > 0, \Delta S > 0, w = 0$

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

5. $\Delta H > 0, \Delta S > 0, w < 0$ **correct**

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 and a negative change in work (work is done on the system via compression).

002   5.0 points

A 23 mL sample of liquid water at 39°C is frozen and cooled to a final temperature of -18°C. Calculate the heat of this process.

1. 11440 J

2. 12300 J

3. -12300 J **correct**

4. -8547 J

5. -534.8 J

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

1) Cool liquid water from 39°C to 0°C:

$$q = mC\Delta T = (23\,\text{g})(4.184\,\frac{\text{J}}{\text{g°C}})[0°C - (39\,°C)]$$
$$= -3753\,\text{J}$$

2) Freeze the ice at constant temperature:

$$q = (m)(-\Delta H_{\text{fus}}) = (23\,\text{g})(334\,\frac{\text{J}}{\text{g}})$$
$$= -7682\,\text{J}$$

3) Cool the ice from 0°C to -18°C:

$$q = mC\Delta T = (23\,\text{g})(2.09\,\frac{\text{J}}{\text{g°C}})[-18°C - 0°C]$$
$$= -865.3\,\text{J}$$

Total heat required:

$(-3753\,\text{J}) + (-7682\,\text{J}) + (-865.3\,\text{J}) = -12300\,\text{J}$

003   5.0 points

For which of the following reactions is $\Delta U_{\text{sys}} > \Delta H_{\text{sys}}$ at constant external pressure?

1. $SnO_2(g) + C(s) \rightarrow Sn(s) + CO_2(g)$

2. $N_2(g) + O_2(g) \rightarrow 2NO(g)$

3. $2NH_3(g) \rightarrow 3H_2(g) + N_2(g)$

4. $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$ **correct**
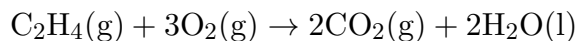
5. $N_2O_5(l) + H_2O(l) \rightarrow 2HNO_3(aq)$

Explanation:
Based on the first law equation:

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

it can be seen that this question is essentially asking to either identify the compression reaction ($\Delta U_{\text{sys}} > \Delta H_{\text{sys}}$) or the expansion reaction ($\Delta U_{\text{sys}} < \Delta H_{\text{sys}}$).

The compression reaction:

$$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$$

The expansion reaction:

$$2NH_3(g) \rightarrow 3H_2(g) + N_2(g)$$

---

004   5.0 points
Determine the boiling point of methanol using the following thermodynamic phase data.

$$\Delta H_{\text{vap}} = 38.3 \, \frac{\text{kJ}}{\text{mol}} \qquad \Delta S_{\text{vap}} = 113 \, \frac{\text{J}}{\text{mol} \cdot \text{K}}$$

1. 169 K

2. -339 K

3. 678 K

4. -678 K

5. 0.339 K

6. 339 K **correct**

Explanation:
   A phase change is in physical equilibrium where $\Delta G = 0$. Therefore, you can rearrange the free energy equation so that:

$T = \Delta H / \Delta S$

$T = (38.3 \, \frac{\text{kJ}}{\text{mol}} \times 1000 \, \frac{\text{J}}{\text{kJ}}) / 113 \, \frac{\text{J}}{\text{mol} \cdot \text{K}}$

$T = 339 \, \text{K}$

---

005   5.0 points
A chemical reaction releases 86.9 kJ heat while 27.7 kJ work is done by the system via expansion. What is the change in the internal energy (ΔU) for the system?

1. -120 kJ

2. 59.2 kJ

3. -64.2 kJ

4. -59.2 kJ

5. -115 kJ **correct**

6. -142 kJ

Explanation:

$$\Delta U_{\text{sys}} = q + w$$
$$-115\text{kJ} = (-86.9\text{kJ}) + (-27.7\text{kJ})$$

---

006   5.0 points
A theoretical reaction has the following thermodynamic data:

$$\Delta H^\circ = 11 \, \text{kJ/mol} \qquad \Delta S^\circ = 87 \, \text{J/mol}$$

   What is the ΔG° for this reaction? Is this reaction spontaneous or non-spontaneous? Assume this reaction is run at 25°C.

1. -14.9 kJ, non-spontaneous

2. 8.82 kJ, non-spontaneous

3. -14.9 kJ, spontaneous **correct**

4. -9.43 kJ, spontaneous

5. -25920 kJ, spontaneous

6. -14930 kJ, non-spontaneous

Explanation:
Using the free energy equation:

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

$$-14.9 = 11 - (298)(87/1000)$$

If ΔG = -14.9 kJ, the reaction is spontaneous.

---

007   5.0 points
You have equal masses of each of the following substances at room temperature:

$$C_{\text{lithium}} = 3.58 \, \frac{\text{J}}{\text{g}^\circ\text{C}} \qquad C_{\text{iron}} = 0.450 \, \frac{\text{J}}{\text{g}^\circ\text{C}}$$

$$C_{\text{air}} = 1.012 \, \frac{\text{J}}{\text{g}^\circ\text{C}} \qquad C_{\text{helium}} = 5.193 \, \frac{\text{J}}{\text{g}^\circ\text{C}}$$

$$C_{\text{mercury}} = 0.140 \, \frac{\text{J}}{\text{g}^\circ\text{C}}$$

Assuming no phase change occurs, which substance will have the highest final temperature after 22.7 J heat is added to each?

1. Iron

2. Lithium

3. The temperature change will be equal for each substance

4. Helium

5. Air

6. Mercury **correct**

Explanation:
Rearranging the heat equation:

$$q = mC_s \Delta T$$

You can see that C_S is inversely proportional to ΔT. Therefore, the substance with the highest heat capacity will have the lowest change in temperature. This is helium, because helium resists change in temperature the most. On the opposite side, mercury has the lowest heat capacity and will therefore have the highest change in temperature.

---

008   5.0 points
The spontaneous vaporization of nitrogen at room temperature is an endothermic reaction. Which of the following statements is true about this process?

1. None of the statements are correct

2. There is a positive ΔS for the system $(\Delta S_{sys} > 0)$ **correct**

3. The magnitude of ΔS for the surroundings must be greater than the magnitude of ΔS for the system $(|\Delta S_{surr}| > |\Delta S_{sys}|)$

4. There is no ΔS for the universe $(\Delta S_{universe} = 0)$

5. There is a negative ΔS for the universe $(\Delta S_{universe} < 0)$

Explanation:
According to the second law of thermodynamics, a spontaneous process has a positive change in entropy for the universe. An exothermic change has a negative change in entropy for the system and a positive change in entropy for the surroundings. In order for this reaction to be spontaneous, the magnitude of ΔS for the surroundings has to be greater than the magnitude of ΔS for the system. An endothermic change has a positive change in entropy for the system and a negative change in entropy for the surroundings.

---

009   5.0 points
For which of the following chemical equations would $\Delta H^\circ_{rxn} = \Delta H^\circ_f$?

1. $2C(s, graphite) + O_2(\ell) \rightarrow 2CO(g)$

2. $N_2(\ell) + 3Cl_2(g) \rightarrow 2NCl_3(\ell)$

3. $\frac{1}{2}N_2(g) + \frac{3}{2}H_2(g) \rightarrow NH_3(g)$ **correct**

4. $CH_4(g) \rightarrow C(s, graphite) + 2H_2(g)$

5. $CH_4(g) + 2Cl_2(g) \rightarrow CCl_4(\ell) + 2H_2(g)$

Explanation:
If $\Delta H^\circ_{rxn} = \Delta H^\circ_f$, the reaction written must be a formation reaction. A formation reaction is a single mole of products formed from its constituent elements in their standard states.

---

010   5.0 points
Unlike state functions such as temperature, volume, and pressure, we consider enthalpy to be a "made-up" state function. Why is enthalpy so important?

1. The change in enthalpy (ΔH) is a measurement of heat at constant pressure and sets an important reference point for reactions in our atmosphere **correct**

2. The sign of the change in enthalpy (ΔH) is a direct indication of the spontaneity of a reaction

3. The change in enthalpy (ΔH) is a measurement of heat at constant pressure, which is directly measured by constant-volume gas phase reactions

4. The change in enthalpy (ΔH) is always equal to the sum of all kinetic and potential energy in a chemical system

5. Enthalpy is a measurement of heat at constant volume, which is an important standard for combustion reactions

Explanation:
The change in enthalpy (ΔH) is a measurement of heat at constant pressure and sets an important reference point for reactions in our atmosphere.

---

011   5.0 points
Which of the following will NOT result in a positive change in entropy for the system?

1. $MgCl_2$ is dissolved in water

2. Water is heated from room temperature to just below its boiling point

3. Sweat evaporates on your skin

4. Two gases are mixed together in a container

5. Half of the gas particles are removed from a container **correct**

Explanation:
Removing particles from a system will lower the total energy content and reduce the total entropy of the system.

---

012   5.0 points
Calculate the ΔS_{surr} for the following reaction at 10°C and 1 atm.
$$CH_3F(\ell) \rightarrow CH_3F(g) \quad \Delta H^\circ_{rxn} = +16.9 \text{ kJ}$$
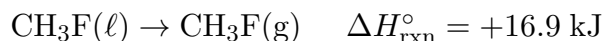
1. -59.7 J/K **correct**

2. -1.69 J/K

3. 1690 J/K

4. 1.69 J/K

5. -845 J/K

6. 59.7 J/K

Explanation:
In general for any process:
$$\Delta S_{surr} = \frac{-\Delta H_{sys}}{T_{surr}}$$

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

therefore $\Delta S_{surr} = -16900/(10 + 273.15)$
$$= -59.7 \text{ J/K}$$

---

013   5.0 points
Consider a reaction that has a negative change in entropy and a positive change in enthalpy. Which of the following conditions will favor spontaneity?

1. High Temperatures

2. Low Temperatures

3. This reaction is non-spontaneous at all temperatures **correct**

4. This reaction is spontaneous at all temperatures

Explanation:
For this reaction, both the entropy and enthalpy terms are unfavorable. This reaction will never be spontaneous

---

014   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:

$$2NO_2(g) + NaCl(s) \rightarrow NOCl(g) + NaNO_3(s)$$
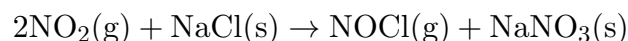
If two moles of NO_2 gas are reacted with excess NaCl to completion at 29°C, what is the value of work for this reaction? Is work done on or by the system?

1. -5024 J, by the system

2. 2512 J, on the system **correct**

3. 2512 J, by the system

4. -2512 J, on the system

5. -5024 J, on the system

6. -241.1 J, by the system

Explanation:
$$w = -\Delta nRT$$

$$w = -(-1\text{mol})(8.314\text{J/molK})(29°\,\text{C}+273.15)$$

$$= 2512\,\text{J}$$

Due to the negative change in gas moles in the piston, work is done on the system via compression.

---
015   5.0 points
Which of the following statements is/are true regarding the first law of thermodynamics?

  I. The internal energy change of the system is equal to the heat plus the work

 II. $\Delta U_{\text{sys}} = -\Delta U_{\text{surr}}$

III. A positive change in internal energy for the system must be matched with a negative change in internal energy for the universe

 IV. $\Delta U_{\text{universe}} = \Delta U_{\text{sys}} + \Delta U_{\text{surr}}$

1. I, II, III and IV

2. I and IV only

3. I, II, and IV only **correct**

4. I only

5. II only

6. III and IV only

7. I and II only

Explanation:
The first law of thermodynamics can be summed up with the idea that energy is always conserved in the universe. Mathematically, this means that:

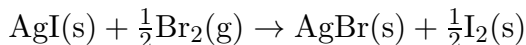$$\Delta U_{\text{sys}} = -\Delta U_{\text{surr}}$$

$$\Delta U_{\text{universe}} = \Delta U_{\text{sys}} + \Delta U_{\text{surr}}$$

$$\Delta U_{\text{universe}} = 0$$

Internal energy is defined as the total energy content of a system. This can be represented as the sum of heat and work or as the sum of potential and kinetic energy.

---
016   5.0 points
Consider the following reaction and standard heats of formation:

$AgI(s) + \frac{1}{2}Br_2(g) \rightarrow AgBr(s) + \frac{1}{2}I_2(s)$

  ΔH_f° for $AgI(s)$ = -61.9 kJ/mol

  ΔH_f° for $AgBr(s)$ = -100.4 kJ/mol

  ΔH_f° for $Br_2(g)$ = +30.9 kJ/mol

What is the $\Delta H_{\text{rxn}}^\circ$?

1. -146.9 kJ/mol

2. $+69.4$ kJ/mol

3. -54.0 kJ/mol **correct**

4. -69.4 kJ/mol

5. -61.9 kJ/mol

Explanation:
$\Delta H_{\text{rxn}} = \Sigma n\Delta H_{\text{f-prod}} - \Sigma n\Delta H_{\text{f-react}}$

$$\Delta H_{\text{rxn}} = [-100.4\,\frac{\text{kJ}}{\text{mol}} + \frac{1}{2}(0\,\frac{\text{kJ}}{\text{mol}})]$$

$$- [-61.9\text{kJ/mol} + \frac{1}{2}(30.9\,\frac{\text{kJ}}{\text{mol}})]$$

$\Delta H_{\text{rxn}} = -54.0$ kJ/mol

---
017   5.0 points

Calculate the value of work when a gas expands from 13.5 to 22.9 liters against a pressure of 3.4 atm at constant temperature. Is work done on or by the system?

1. +0.3154 J, on the system

2. -0.3154 J, by the system

3. -3238 J, by the system **correct**

4. -3238 J, on the system

5. +3238 J, on the system

6. -0.3154 J, on the system

Explanation:

$$\text{work} = -P\Delta V = -(3.4\,\text{atm})(22.9\,\text{L} - 13.5\,\text{L})$$
$$= -31.96\,\text{L}\cdot\text{atm}$$

$$\text{work} = -31.96\,\text{L}\cdot\text{atm} \times \tfrac{101.325\,\text{J}}{1\,\text{L}\cdot\text{atm}} = -3238\,\text{J}$$

---

018   5.0 points

The unknown element X is a metallic solid that oxidizes at room temperature and pressure to form X_2O_3. Given the following thermodynamic data, calculate the ΔG_f°.
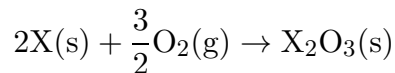
Substance	ΔH_f° kJ/ mol	S° J/ mol K
X(s)	–	38.3
O_2(g)	–	205
O(g)	249	161
X_2O_3(s)	−860	92.2

1. 6.09×10^6 kJ mol^{-1}

2. 1470 kJ mol^{-1}

3. 742.35 kJ mol^{-1}

4. -773 kJ mol^{-1} **correct**

5. −5210 kJ mol^{-1}

6. −1470 kJ mol^{-1}

7. 773 kJ mol^{-1}

8. −742.35 kJ mol^{-1}

Explanation:

For the formation of one mole of X_2O_3, you can use the chemical equation shown below:

$$2X(s) + \frac{3}{2}O_2(g) \rightarrow X_2O_3(s)$$

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

Using ΔH_f° and S° values:

$$\Delta S = 92.2 - [(1.5 \times 205) + (2 \times 38.3)]$$

$$\Delta S = -291.9\text{J/K}$$

$$\Delta H = -860\text{kJ}$$

$$\Delta G = -773\text{kJ}$$

This was solved per mole based on the formation reaction.

---

019   5.0 points

When a 0.401 g sample of a clean-burning hydrocarbon (molecular weight = 42.05 g/mol) is combusted in a rigid container, the temperature increases from 24.987°C to 28.623°C. The total volume of water is 0.746 L. The sum of all hardware components of the calorimeter have a heat capacity of 2.55 kJ/°C. Calculate the internal energy of combustion for one mole of this hydrocarbon.

1. -3.71799 kJ/mol

2. +9277.90 kJ/mol

3. +51.4232 kJ/mol

4. +2162.35 kJ/mol

5. +3.71799 kJ/mol

6. -2162.35 kJ/mol **correct**

7. +20.6207 kJ/mol

8. -20.6207 kJ/mol

Explanation:
Heat released by the combustion reaction:

$$q = mC\Delta T + C\Delta T$$

$$q_{cal} = (746)(4.184)(3.636) + (2.55 \times 1000)(3.636)$$

$$q_{cal} = 20620.7 J$$

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

$$q_v = -20.6207 kJ$$

This is how much heat was released when 0.401 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.
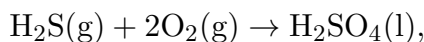
$$\frac{-20.6207\,\text{kJ}}{0.401\,\text{g}} \times \frac{42.05\,\text{g}}{1\,\text{mol}} = -2162.35\,\text{kJ/mol}$$
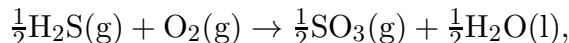

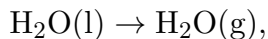**020   5.0 points**
Consider the following balanced chemical reaction:

$$H_2SO_4(l) \rightarrow SO_3(g) + H_2O(g)$$

To solve for the ΔH°_{rxn} for this reaction, you collect the following data in the lab:

$H_2S(g) + 2O_2(g) \rightarrow H_2SO_4(l)$,

$$\Delta H^\circ_{rxn} = \text{-232.3 kJ/mol}$$

$\frac{1}{2}H_2S(g) + O_2(g) \rightarrow \frac{1}{2}SO_3(g) + \frac{1}{2}H_2O(l)$,

$$\Delta H^\circ_{rxn} = \text{-102.8 kJ/mol}$$

$H_2O(l) \rightarrow H_2O(g)$,

$$\Delta H^\circ_{rxn} = 51.9 \text{ kJ/mol}$$

What is the ΔH°_{rxn} of the overall balanced reaction?

1. -386 kJ/mol

2. 490 kJ/mol

3. 181 kJ/mol

4. -516 kJ/mol

5. 78.6 kJ/mol **correct**

Explanation:
The objective is to arrange the three reactions so that the sum of the reactions gives you the overall balanced chemical reaction. You will need to make two modifications. The first reaction must be reversed, which means you will flip the sign of the enthalpy of reaction. The coefficients of the second reaction must be doubled, which means you will need to multiple the enthalpy of reaction by two. From there:

$\Delta H^\circ_{rxn} = -(-232.3) + (2 \times -102.8) + (51.9) = 78.6 kJ/mol$