

5 – Thermochemistry/Thermodynamics

WS#1: WORK/ENERGY/ENTHALPY

- Which has the greater kinetic energy, an object with a mass of 2.0kg and a velocity of 1.0m/s or an object with a mass of 1.0kg and a velocity of 2.0m/s? (Ans: $1.0\text{kg}\cdot\text{m}^2/\text{s}^2$ (J) vs. $2.0\text{kg}\cdot\text{m}^2/\text{s}^2$ (J))
- 900.J of heat are added to a system and 200. joules of work are done on the system. What is the change in energy of the system? (Ans: +1100J)
- Calculate the work for a system that absorbs 260kJ of heat and for which the change in energy is +157kJ. Is the work done on or by the system? Does the system expand or contract? (Ans: -103kJ)
- Two liters of an ideal gas is expanded into a vacuum and the new volume is 4L. If 500J of heat flow out of the container, what is the change in internal energy of the system? (Ans: -500J)
- One of the components of polluted air is NO. It is formed in the high temperature environment of internal combustion engines by the following reaction: $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g})$ ($\Delta H = 180\text{kJ}$). Why are high temperatures needed to convert N_2 and O_2 to NO?
- Are the following processes exothermic or endothermic?
 - When solid KBr is dissolved in water, the solution gets colder.
 - Natural gas, CH_4 , is burned in a furnace.
 - When concentrated H_2SO_4 is added to water, the solution gets very hot.
 - Water is boiled in a tea kettle.
- The overall reaction in commercial heat packs can be represented as $4\text{Fe}(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{Fe}_2\text{O}_3(\text{s})$ ($\Delta H = -1652\text{kJ}$)
 - How much heat is released when 4.00mol iron is reacted with excess O_2 ? (Ans: -1652kJ)
 - How much heat is released when 1.00mol of Fe_2O_3 is produced? (Ans: -826kJ)
 - How much heat is released when 1.00g of iron is reacted with excess O_2 ? (Ans: -7.40kJ)
 - How much heat is released when 10.0g Fe and 2.00g O_2 are reacted? (Ans: -34.4kJ)
- How much heat is transferred when 147g of NO_2 reacts with 100.g H_2O according to the following equation:

$$3\text{NO}_2 + \text{H}_2\text{O} \rightarrow 2\text{HNO}_3 + \text{NO} \quad \Delta H = -138 \text{ kJ} \text{ (Ans: -147kJ)}$$

WS#2: CALORIMETRY

Measuring heat (formerly measured in calories) is called **calorimetry**. Now we measure heat energy in Joules (J). The equation we use is:

$$q = m \times C_p \times \Delta T$$

q = heat energy, m = mass of water, C_p = the specific heat capacity, and
 ΔT = the change in temperature (in $^\circ\text{C}$ or K)

- Water has a specific heat capacity of 4.184 J/g \cdot $^\circ\text{C}$. This means it takes 4.184J to heat 1.00 gram of water 1.00 $^\circ\text{C}$.
 - How much energy will it take to heat 10.0 grams of water 1 $^\circ\text{C}$? _____ (Ans: 41.84J)
 - How much energy is needed to heat 30.0g H_2O from 10.0 $^\circ\text{C}$ to 50.0 $^\circ\text{C}$? _____ (Ans: 5020.8J)
- Let's try a standard calorimetry problem. A pot of water (2.5 liters of water) initially at 25.0 $^\circ\text{C}$ is heated to boiling (100. $^\circ\text{C}$).
 - How much energy (in J) is needed to heat the water? (The density of water is 1.00g/mL.). (Ans: 780,000J)
 - What would this amount of heat be in kJ? _____ (Ans: 780kJ)
- What amount of heat is *released* when 175g of water *cools* from 100. $^\circ\text{C}$ to room temperature, 20.0 $^\circ\text{C}$? (Ans: -59kJ)
- We don't always have to warm up or cool down water. The specific heat capacity of copper metal is 0.39J/g \cdot $^\circ\text{C}$.
 - It is _____ (easier/more difficult) to heat up copper than to heat up water.
 - How much energy would it take to heat up a 5.20g sample of copper from 20.0 $^\circ\text{C}$ to 100. $^\circ\text{C}$? (Ans: 160J)
- If 300.J of heat energy were used to heat up a 5.00 gram sample of copper metal and a 5.00 gram sample of water both starting at 10.0 $^\circ\text{C}$, calculate the final temperature of each sample? (Ans: 163.8 $^\circ\text{C}$ and 24.3 $^\circ\text{C}$)

Signs of ΔT and q : $-q$ means heat is **released**. $+q$ means heat is **absorbed**. ΔT is always **final** temperature – **initial** temperature ($T_f - T_i$). If something is getting **hotter** (10 $^\circ \rightarrow$ 30 $^\circ$), the ΔT is 30 – 10 = + 20 $^\circ$. (heat is **absorbed**)
 If something is getting **cooler** (75 $^\circ \rightarrow$ 25 $^\circ$), the ΔT is 25 – 75 = – 50 $^\circ$. (heat is **released**)

- Suppose we mix 90.0 grams of hot water (90.0 $^\circ\text{C}$) with 10.0 grams of cold water (10.0 $^\circ\text{C}$). Let x = the final temperature. $C = 4.184 \text{ J/g}\cdot^\circ\text{C}$
 - Set up an expression for the energy *released* (q) by the hot water ($\Delta q_{\text{hot}} = m_{\text{hot}} \times C_p \times \Delta T_{\text{hot}}$)
 - Set up an expression for the energy *absorbed* (q) by the cold water ($\Delta q_{\text{cold}} = m_{\text{cold}} \times C_p \times \Delta T_{\text{cold}}$)

- c. Knowing that the heat released = – heat absorbed, combine the two expressions and solve for x . (Ans: 82°C)
7. We don't always have to use water. Let's use some aluminum shot. 175 grams of hot aluminum (100.°C) is dropped into an insulated cup that contains 40.0 mL of ice cold water (0.0°C). Follow the example above to determine the final temperature, x . (a) Set up an expression for the heat lost by the aluminum ($C_p = 0.900 \text{ J/g}\cdot^\circ\text{C}$). (b) Set up an expression for the heat gained by the cold water. (c) Put the two expressions together (don't forget to change one of the signs) and solve for x . (Ans: 48.5°C)
8. Somewhat Confusing Definitions:
 There are several terms used in this chapter that sound very similar. Use the data provided to calculate each of them to clarify the differences. I've added some "Notes" that I hope will help. 74.8 J of heat is required to raise the temperature of 18.69 g of silver from 10.0°C to 27.0°C.
- What is the heat capacity of the silver sample? ($\text{J}/^\circ\text{C}$). Note: This is a useful value only for this specific sample of silver. (Ans: 4.4 J/°C)
 - What is the specific heat capacity of silver? ($\text{J}/\text{g}\cdot^\circ\text{C}$) (Ans: 0.235 J/g·°C)
- Note: This is a useful value for *any* sample of silver that is heated or cooled.
 This is equivalent to the $4.184 \text{ J}\cdot\text{g}^{-1}\cdot^\circ\text{C}^{-1}$ that we use for water. This value is also called the specific heat.
9. 1.150 g of sucrose goes through combustion in a bomb calorimeter. If the temperature rose from 23.42°C to 27.64°C and the heat capacity of the calorimeter is 4.90 kJ/°C, then determine the heat of combustion of sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$, in kJ per mole of $\text{C}_{12}\text{H}_{22}\text{O}_{11}$. (Ans: -6160kJ/mol)
10. The following data was collected in an experiment similar to the Specific Heat experiment performed in class. Fill in the missing values. (Assume the calorimeter has a calorimeter constant of 0 J/°C)

Data & Calculations		Glass Beads
mass of glass beads		4.88 g
mass of water		10.14 g
initial temperature of water		19.2 °C
initial temperature of beads		89.2 °C
final temperature of mixture		24.8 °C
temp change of water (°C)		
temp change of hot beads (°C)		
change in energy of water, q_{water} (J)		
change in energy of calorimeter (J)		0 J
change in energy of beads, q_{beads} (J)		
specific heat of beads ($\text{J}\cdot\text{g}^{-1}\cdot^\circ\text{C}^{-1}$)		
accepted value of specific heat		.833
% error		

WS#3: HEAT OF FORMATION (ΔH_f°), HESS'S LAW, AND BOND ENTHALPY

1. Calculate the standard enthalpy of combustion for the following reaction: $\text{C}_6\text{H}_{12}\text{O}_6(\text{s}) + 6\text{O}_2(\text{g}) \rightarrow 6\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l})$. To solve this problem, we must know the following ΔH_f° values: (Ans: -2801kJ/mol)

$\text{C}_6\text{H}_{12}\text{O}_6(\text{s})$	-1275.0	$\text{CO}_2(\text{g})$	-393.5
$\text{O}_2(\text{g})$	zero	$\text{H}_2\text{O}(\text{l})$	-285.8

2. Complete combustion of 1.00 mol of acetone ($\text{C}_3\text{H}_6\text{O}$) liberates 1790 kJ: $\text{C}_3\text{H}_6\text{O}(\text{l}) + 4\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l}) + 1790 \text{ kJ}$. Using this information along with the data below (values in kJ/mol), calculate the enthalpy of formation of acetone. ($\Delta H_f^\circ \text{CO}_2$: -393.5kJ/mol and $\Delta H_f^\circ \text{H}_2\text{O}$: -285.83kJ/mol) (Ans: -248kJ/mol)
3. Ammonia reacts with oxygen to form nitrogen dioxide and steam, as follows: $4\text{NH}_3(\text{g}) + 7\text{O}_2(\text{g}) \rightarrow 4\text{NO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{g})$. Given the following standard enthalpies of formation (given in kJ/mol), calculate the enthalpy of the reaction. Note that water is given as a gas. Normally, water as a liquid is used in problems. Not in this one.
 (Ans: -1135kJ)

$\text{NH}_3(\text{g})$	-45.90	$\text{NO}_2(\text{g})$	+33.1	$\text{H}_2\text{O}(\text{g})$	-241.83
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4. The standard enthalpy change, ΔH° , for the thermal decomposition of silver nitrate according to the following equation is +78.67 kJ: $\text{AgNO}_3(s) \rightarrow \text{AgNO}_2(s) + (1/2)\text{O}_2(g)$. The standard enthalpy of formation of $\text{AgNO}_3(s)$ is -123.02 kJ/mol. Calculate the standard enthalpy of formation of $\text{AgNO}_2(s)$ (Ans: -44.35kJ)
5. Calculate the enthalpy for this reaction: $2\text{C}(s) + \text{H}_2(g) \rightarrow \text{C}_2\text{H}_2(g)$. (Ans: +226.7kJ)
- $$2\text{C}_2\text{H}_2(g) + 5\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 2\text{H}_2\text{O}(l) \quad \Delta H^\circ = -2599 \text{ kJ}$$
- $$\text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g) \quad \Delta H^\circ = -393.5 \text{ kJ}$$
- $$2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l) \quad \Delta H^\circ = -571.6 \text{ kJ}$$
6. Given the following information:

$2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)$	$\Delta H^\circ = -116 \text{ kJ}$
$2\text{N}_2(g) + 5\text{O}_2(g) + 2\text{H}_2\text{O}(l) \rightarrow 4\text{HNO}_3(aq)$	$\Delta H^\circ = -256 \text{ kJ}$
$\text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g)$	$\Delta H^\circ = +183 \text{ kJ}$

- Calculate the enthalpy change for the reaction below: $3\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow 2\text{HNO}_3(aq) + \text{NO}(g)$ (Ans: -137kJ)
7. Calculate ΔH for this reaction: $\text{ClF}(g) + \text{F}_2(g) \rightarrow \text{ClF}_3(g)$ given: (Ans: -108.7kJ)

$2\text{ClF}(g) + \text{O}_2(g) \rightarrow \text{Cl}_2\text{O}(g) + \text{F}_2\text{O}(g)$	$\Delta H^\circ = 167.4 \text{ kJ}$
$2\text{ClF}_3(g) + 2\text{O}_2(g) \rightarrow \text{Cl}_2\text{O}(g) + 3\text{F}_2\text{O}(g)$	$\Delta H^\circ = 341.4 \text{ kJ}$
$2\text{F}_2(g) + \text{O}_2(g) \rightarrow 2\text{F}_2\text{O}(g)$	$\Delta H^\circ = -43.4 \text{ kJ}$

8. Using bond enthalpies, calculate the reaction enthalpy (ΔH) for: $\text{CH}_4(g) + \text{Cl}_2(g) \rightarrow \text{CH}_3\text{Cl}(g) + \text{HCl}(g)$. Bond enthalpies (in kJ/mol): C-H (413); Cl-Cl (239); C-Cl (339); H-Cl (427) (Ans: -114kJ)
9. What is the enthalpy of reaction for the following equation: $2 \text{CH}_3\text{OH}(l) + 3 \text{O}_2(g) \rightarrow 2 \text{CO}_2(g) + 4 \text{H}_2\text{O}(g)$. Given the following bond enthalpies (in kJ/mol): C-H (414); C-O (360); C=O (799); O=O (498); O-H (464) (Ans: -1282kJ)
10. Calculate the bond energy of the Cl-F bond using the following data: $\text{Cl}_2(g) + \text{F}_2(g) \rightarrow 2\text{ClF}(g)$, $\Delta H = -108 \text{ kJ}$. Bond enthalpies (in kJ/mol): Cl-Cl (239); F-F (159) (Ans: 253kJ for the Cl-F bond)
11. The reaction of H_2 with F_2 produces HF with $\Delta H = -269 \text{ kJ/mol}$ of HF. If the H-H and H-F bond energies are 432 and 565 kJ/mol, respectively, what is the F-F bond energy? $\text{H}_2(g) + \text{F}_2(g) \rightarrow 2\text{HF}(g)$ (Ans: 160kJ for the F-F bond)

WS#4: Entropy

- Which of the following processes are spontaneous?
 - Salt dissolves in water.
 - A clear solution becomes a uniform color after a few drops of dye are added.
 - Iron rusts
 - You clean your bedroom.
- Choose the compound with the greatest positional entropy in each case.
 - 1mol H_2 (at STP) or 1mol H_2 (at 100°C, 0.5atm)
 - 1mol N_2 (at STP) or 1mol N_2 (at 100K, 2.0atm)
 - 1mol $\text{H}_2\text{O}(s)$ (at 0°C) or 1mol $\text{H}_2\text{O}(l)$ (at 20°C)
- Which of the following involves an increase in entropy of the system?
 - melting of a solid
 - sublimation
 - freezing
 - mixing
 - separation
 - boiling
- Predict the sign of ΔS_{surr} for the following processes.
 - $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g)$
 - $\text{CO}_2(g) \rightarrow \text{CO}_2(s)$
 - $2\text{H}_2\text{S}(g) + \text{SO}_2(g) \rightarrow 3\text{S}_{\text{rhombic}}(s) + 2\text{H}_2\text{O}(g)$
 - $2\text{SO}_3(g) \rightarrow 2\text{SO}_2(g) + \text{O}_2(g)$
 - $\text{Fe}_2\text{O}_3(s) + 3\text{H}_2(g) \rightarrow 2\text{Fe}(s) + 3\text{H}_2\text{O}(g)$
- Calculate ΔS_{surr} for the following reactions at 25°C and 1atm.
 - $\text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g)$ $\Delta H^\circ = -2221\text{kJ}$ (Ans: +7453J/K)
 - $2\text{NO}_2(g) \rightarrow 2\text{NO}(g) + \text{O}_2(g)$ $\Delta H^\circ = 112\text{kJ}$ (Ans: -376J/K)
- Ethanethiol ($\text{C}_2\text{H}_5\text{SH}$; also called ethyl mercaptan) is commonly added to natural gas to provide the "rotten egg" smell of a gas leak. The boiling point of ethanethiol is 35°C and its heat of vaporization is 27.5kJ/mol. What is the entropy of vaporization for this substance? (An:89.3J/K-mol)
- For ammonia, NH_3 , the enthalpy of fusion is 5.65kJ/mol and the entropy of fusion is 28.9J/K-mol.
 - Will $\text{NH}_3(s)$ spontaneously melt at 200.K?
 - What is the approximate melting point of ammonia? (Ans: a. yes, $\Delta G < 0$, b. 196K)
- Determine the standard entropy change for the decomposition of one mole of solid calcium carbonate, forming solid calcium oxide and carbon dioxide gas. $\text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g)$ (ΔS° for $\text{CaO}(s)$ is 39.8 J/mol-K, ΔS° for $\text{CO}_2(g)$ is 213.7 J/mol-K, and ΔS° for $\text{CaCO}_3(s)$ is 92.9 J/mol-K) (Ans: 160.6J/mol-K)

WS#5: Free Energy

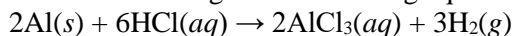
- Define Gibb's Free energy.
 - Given the values for ΔH and ΔS , which of the following changes will be spontaneous at constant T and P ?
 - $\Delta H = +25\text{kJ}$, $\Delta S = +5.0\text{J/K}$, $T = 300\text{K}$
(Ans: 23.5kJ/mol)
 - $\Delta H = +25\text{kJ}$, $\Delta S = +100\text{J/K}$, $T = 300\text{K}$
(Ans: -5kJ/mol)
 - $\Delta H = -10\text{kJ}$, $\Delta S = +5.0\text{J/K}$, $T = 298\text{K}$
(Ans: -11.5kJ/mol)
 - $\Delta H = -10\text{kJ}$, $\Delta S = -40\text{J/K}$, $T = 200\text{K}$
(Ans: -2.0kJ/mol)
 - Describe how the sign of ΔG determines the spontaneity of a reaction:
 - $\Delta G > 0$
 - $\Delta G < 0$
 - $\Delta G = 0$
 - Given the following thermodynamics values, determine how temperature affects the spontaneity of the reaction. Is the reaction enthalpy, entropy driven, both, or neither (compare ΔH vs $(-T\Delta S)$).
 - $\Delta H = +327\text{ kJ/mol}$, $\Delta S = +127\text{ J/K-mol}$
 - $\Delta H = -187\text{ kJ/mol}$, $\Delta S = -298\text{ J/K-mol}$
 - $\Delta H = +194\text{ kJ/mol}$, $\Delta S = -4.0\text{ J/K-mol}$
 - $\Delta H = -397\text{ kJ/mol}$, $\Delta S = +653\text{ J/K-mol}$
 - For each reaction below, determine the temperature at which the reaction will be spontaneous.
 - $\Delta H = +245\text{ kJ/mol}$, $\Delta S = +48\text{ J/K-mol}$
(Ans: 5104K)
 - $\Delta H = +187\text{ kJ/mol}$, $\Delta S = +365\text{ J/K-mol}$
(Ans: 512K)
 - $\Delta H = -456\text{ kJ/mol}$, $\Delta S = -38\text{ J/K-mol}$
(Ans: 12,000K)
 - $\Delta H = -547\text{ kJ/mol}$, $\Delta S = -97\text{ J/K-mol}$
(Ans: 5639K)
 - $\Delta H^\circ = 131.4\text{ kJ}$ and $\Delta G^\circ = 91.2\text{ kJ}$ for the reaction $\text{C}(s) + \text{H}_2\text{O}(g) \rightarrow \text{CO}_2(g) + \text{H}_2(g)$. Calculate ΔS° in J/K for this reaction at 25°C . (Ans: 135 J/K)
 - Using the values of ΔG°_f , calculate the $\Delta G^\circ_{\text{rxn}}$ of each of the reactions below. Is each reaction spontaneous at standard conditions? $\Delta G^\circ_f(\text{N}_2\text{O}_4) = 100\text{ kJ/mol}$, $\Delta G^\circ_f(\text{NO}_2) = 51\text{ kJ/mol}$, $\Delta G^\circ_f(\text{NH}_4\text{Cl}) = -203\text{ kJ/mol}$, $\Delta G^\circ_f(\text{NH}_3) = -16\text{ kJ/mol}$, $\Delta G^\circ_f(\text{HCl}) = -95\text{ kJ/mol}$, $\Delta G^\circ_f(\text{Fe}_2\text{O}_3) = -742\text{ kJ/mol}$, $\Delta G^\circ_f(\text{H}_2\text{O}) = -229\text{ kJ/mol}$
 - $\text{N}_2\text{O}_4(g) \rightarrow 2\text{NO}_2(g)$ (Ans: 2 kJ)
 - $\text{NH}_4\text{Cl}(l) \rightarrow \text{NH}_3(g) + \text{HCl}(g)$ (Ans: 92 kJ)
 - $3\text{H}_2(g) + \text{Fe}_2\text{O}_3(s) \rightarrow 2\text{Fe}(s) + 3\text{H}_2\text{O}(g)$ (Ans: 55kJ)
 - $\text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g)$ (Ans: -32 kJ)
 - Calculate ΔG for the reaction $\text{NO}(g) + \text{O}_3(g) \rightarrow \text{NO}_2(g) + \text{O}_2(g)$ for these conditions:
 $T = 298\text{K}$, $P_{\text{NO}} = 1.00 \times 10^{-6}\text{atm}$, $P_{\text{O}_3} = 2.00 \times 10^{-6}\text{atm}$, $P_{\text{NO}_2} = 1.00 \times 10^{-7}\text{atm}$, and $P_{\text{O}_2} = 1.00 \times 10^{-3}\text{atm}$
 $\Delta G^\circ_f(\text{NO}) = 87\text{kJ/mol}$, $\Delta G^\circ_f(\text{O}_3) = 163\text{kJ/mol}$, and $\Delta G^\circ_f(\text{NO}_2) = 52\text{ kJ/mol}$ (Ans: -188kJ)
 - Coupling of the two reactions below affords the overall reaction: $\text{Fe}_3\text{O}_4(s) + 2\text{C}(s) \rightarrow 3\text{Fe}(s) + 2\text{CO}_2(g)$. What is the value of ΔG° for this reaction? (Ans: +225.4 kJ/mol)
$$\begin{array}{l} \text{Fe}_3\text{O}_4(s) \rightarrow 3\text{Fe}(s) + 2\text{O}_2(g) \quad \Delta G^\circ = 1014.2\text{ kJ} \\ \text{O}_2(g) + \text{C}(s) \rightarrow \text{CO}_2(g) \quad \Delta G^\circ = -394.4\text{ kJ} \end{array}$$
 - Calculate the ΔG° for a reaction for which $K_c = 5.0 \times 10^8$ at 25°C . (Ans: -50 kJ)
 - Given that ΔG_f° for IBr is 1.3 kJ/mol , calculate K_c for the reaction at 25°C : $\text{I}_2 + \text{Br}_2 \leftrightarrow 2\text{IBr}$ (Ans: 0.35)
 - A particular reaction has an equilibrium constant of 0.48 at 25°C . What is the value of ΔG° (in kJ)? (Ans: 1.8)
- (Move to Electrochemistry)
- What is n ?
 - What is F ?
 - Write the equation that relates ΔG° to E° ?
 - For each of the following equations, what must n be?
 - $\text{Cu}^2 + \text{Mg} \rightarrow \text{Mg}^{2+} + \text{Cu}$
 - $2\text{Ag}^{1+} + \text{Sn} \rightarrow 2\text{Ag} + \text{Sn}^{2+}$
 - $2\text{Al} + 3\text{Pb}^{2+} \rightarrow 3\text{Pb} + 2\text{Al}^{3+}$
 - What is the ΔG° of the following batteries?
 - $\text{Cu}^2 + \text{Mg} \rightarrow \text{Mg}^{2+} + \text{Cu}$ (Ans: -523kJ)
 - $2\text{Ag}^{1+} + \text{Sn} \rightarrow 2\text{Ag} + \text{Sn}^{2+}$ (Ans: -181kJ)
 - $2\text{Al} + 3\text{Pb}^{2+} \rightarrow 3\text{Pb} + 2\text{Al}^{3+}$ (Ans: -886kJ)
 - What sign must each of the following be for a battery?
 - E°
 - ΔG°

WS#6: MC Review Questions

Thermochemistry

Use the following information to answer questions 1 and 2.

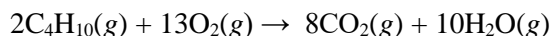
Aluminum reacts with hydrochloric acid according to the following equation:



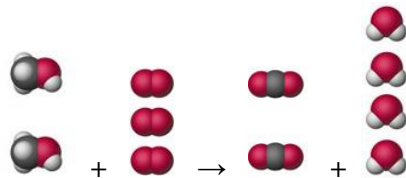
- If a sample of 27.0g of aluminum metal is added to 333mL of 3.0M hydrochloric acid, the volume of hydrogen gas produced at standard temperature and pressure is
 - 2.80L
 - 5.60L
 - 11.2L
 - 22.4L
- What is the approximate density of the hydrogen gas produced at STP?
 - 0.1g/L
 - 0.2g/L
 - 0.3g/L
 - 0.4g/L
- Given the following data.

Calculate the $\Delta H^\circ_{\text{rxn}}$ for the following reaction.

Substance	H_f° (kJ/mol)
$\text{CO}_2(g)$	-393.5
$\text{H}_2\text{O}(g)$	-285.8
$\text{C}_4\text{H}_{10}(g)$	-124.7



- 6255.4kJ
- 5756.6 kJ
- 40.6 kJ
- 539.4 kJ



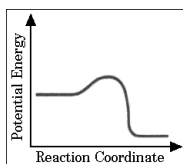
- Consider the reaction shown: $\text{CH}_3\text{OH} + 1.5\text{O}_2 \rightarrow \text{CO}_2 + 1.5\text{H}_2\text{O}$. Given the following information, calculate $\Delta H^\circ_{\text{rxn}}$ for the combustion reaction represented above, where each molecule represents one mole of that substance. Assume that all states are those that are listed in the table

Substance	ΔH_f° (kJ/mol)
$\text{CO}_2(g)$	-393.5
$\text{H}_2\text{O}(g)$	-285.8
$\text{CH}_3\text{OH}(g)$	-201.0

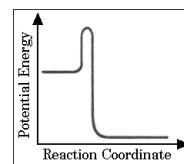
- 2332.2kJ
- 1528.2kJ
- 478.3kJ
- 45.8kJ

- Which of the following graphs describes a pathway of reaction that is exothermic with a high activation energy?

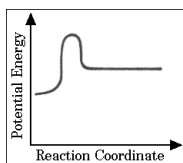
a.



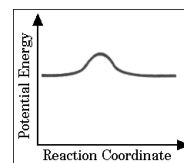
c.



b.



d.



- One cup of ice has a mass of approximately 250g. The ice is at 0.0°C. How much heat is required to melt the ice and then warm the resulting water 25.0°C? A table of useful information is provided.

Heat of fusion of $\text{H}_2\text{O}(s)$	332kJ/kg
Specific heat of $\text{H}_2\text{O}(l)$	4.18J/°C-g

- 25kJ
- 58kJ
- 83kJ
- 110kJ

- A 20.g sample of metal at 100°C is placed into 30.g of kerosene at 35°C. The final temperature of both the kerosene and the metal is 60°C. If the specific heat of kerosene is 2.0J/°C-g, what is the specific heat of the metal?
 - 1.0 J/°C-g
 - 1.5 J/°C-g
 - 2.0 J/°C-g
 - 2.5 J/°C-g

8. Calculate the enthalpy change for the following reaction given the information below. $\text{Al}_2(\text{CO}_3)_3(\text{s}) \rightarrow \text{Al}_2\text{O}_3(\text{s}) + 3\text{CO}_2(\text{g})$.
- $$\begin{array}{ll} \text{C}(\text{graphite}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) & \Delta H = W \text{ kJ} \\ 4\text{Al}(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{Al}_2\text{O}_3(\text{s}) & \Delta H = X \text{ kJ} \\ 2\text{Al}(\text{s}) + 9/2\text{O}_2(\text{g}) + 3\text{C}(\text{graphite}) \rightarrow \text{Al}_2(\text{CO}_3)_3(\text{s}) & \Delta H = Y \text{ kJ} \end{array}$$
- a. $W + X + Y$
 b. $3W + X - Y$
 c. $3W + 1/2X + Y$
 d. $3W + 1/2X - Y$
9. In which of the following isothermal expansions of an ideal gas is the work done by the system the largest? Assume the temperature is the same for all choices.
- a. from 1 to 5L against an opposing pressure of $1/2$ atm.
 b. from 1 to 5L against an opposing pressure of 1 atm.
 c. from 1 to 5L against an opposing pressure of 3 atm.
 d. from 1 to 10L against an opposing pressure of 2 atm.
10. Consider four 100.0g samples of water, each in a separate beaker at 25°C . Into each beaker your drop 10.0g of a different metal that has been heated to 95°C . Assume no heat loss to the surroundings, which water sample will have the lowest final temperature?
- a. the water to which you have added aluminum ($s = 0.89\text{J}/^\circ\text{C}\cdot\text{g}$)
 b. the water to which you have added iron ($s = 0.45\text{J}/^\circ\text{C}\cdot\text{g}$)
 c. the water to which you have added copper ($s = 0.20\text{J}/^\circ\text{C}\cdot\text{g}$)
 d. the water to which you have added lead ($s = 0.14\text{J}/^\circ\text{C}\cdot\text{g}$) $q = 10 \times s \times T$
11. A student performs a neutralization reaction involving an acid and a base in an open polystyrene coffee-cup calorimeter. How would the calculated value of ΔH differ from the actual value if there was significant heat loss to the surroundings?
- a. ΔH_{calc} would be negative, but more negative than actual value.
 b. ΔH_{calc} would be negative, but less negative than actual value.
 c. ΔH_{calc} would be positive, but more positive than actual value.
 d. ΔH_{calc} would be positive, but less positive than the actual value.
12. Which of the following is endothermic?
- a. Water freezes to form ice.
 b. Steam condenses on a bathroom mirror.
 c. Ice cream melts.
 d. Coffee cools as it sits.
13. At 25°C , the following heats of reaction are known:
- $$\begin{array}{ll} 2\text{C}_2\text{H}_2(\text{g}) + 5\text{O}_2 \rightarrow 4\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) & \Delta H = -2600.0\text{kJ} \\ \text{C}(\text{graphite}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) & \Delta H = -394\text{kJ} \\ 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l}) & \Delta H = -574\text{kJ} \end{array}$$
- At the same temperature, calculate the ΔH for the reaction $2\text{C}(\text{graphite}) + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_2(\text{g})$
- a. -2422kJ
 b. -225kJ
 c. 225kJ
 d. 2422kJ
14. Using Hess's law and the equation below find ΔH° at 25°C for the oxidation of $\text{C}_2\text{H}_5\text{OH}(\text{l})$.
- $$\begin{array}{ll} \text{C}_2\text{H}_5\text{OH}(\text{l}) + 3\text{O}_2(\text{g}) \rightarrow 3\text{H}_2\text{O}(\text{l}) + 2\text{CO}_2(\text{g}) & \\ \text{C}_2\text{H}_4(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) & \Delta H^\circ = W \text{ kJ} \\ 2\text{C}(\text{graphite}) + 3\text{H}_2(\text{g}) + 1/2\text{O}_2(\text{g}) \rightarrow \text{C}_2\text{H}_5\text{OH}(\text{l}) & \Delta H^\circ = X \text{ kJ} \\ \text{C}_2\text{H}_4(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{C}_2\text{H}_5\text{OH}(\text{l}) & \Delta H^\circ = Y \text{ kJ} \end{array}$$
- a. $W - Y$
 b. $X - 2Y$
 c. $X + 2W + Y$
 d. $2X - W + Y$
15. A popular demonstration is to drop a piece of sodium metal into water. The products are sodium hydroxide and hydrogen gas. Determine ΔH_{rxn} for this reaction for 1.00mol of hydrogen gas being produced, given $\Delta H^\circ_f [\text{H}_2\text{O}(\text{l})] = -286\text{kJ}/\text{mol}$ and $\Delta H^\circ_f [\text{NaOH}(\text{l})] = -470\text{kJ}/\text{mol}$
- a. -368kJ
 b. -184kJ
 c. 184kJ
 d. 368kJ
16. Consider the freezing of liquid water at -10°C and 1 atm. For this process, what are the signs for ΔH , ΔS , and ΔG ?
- a. $+, -, 0$
 b. $-, +, 0$
 c. $+, +, -$
 d. $-, -, -$
17. A 100mL sample of water is placed in a coffee-cup calorimeter. Solid NaCl is then dissolved in the water. The temperature of the water decreases from 20.5°C to 19.7°C and is then allowed to return to room temperature (20.5°C). Determine the signs for ΔH and ΔS for the process of dissolving NaCl and ΔG for the entire process at constant temperature.
- a. $+, -, 0$
 b. $-, +, 0$
 c. $+, +, -$
 d. $-, -, -$

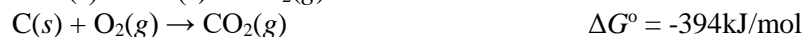
18. Which reaction would have the most positive ΔS ?
- $\text{CO}(g) + 2\text{H}_2(g) \rightarrow \text{CH}_3\text{OH}(l)$
 - $2\text{CH}_3\text{OH}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 4\text{H}_2\text{O}(g)$
 - $\text{HCl}(g) + \text{NH}_3(g) \rightarrow \text{NH}_4\text{Cl}(s)$
 - $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}(s) + 2\text{NH}_3\text{NO}_3(s) \rightarrow \text{Ba}(\text{NO}_3)_2(s) + 2\text{NH}_3(g) + 10\text{H}_2\text{O}(l)$
19. Solutions A and B are both clear and colorless. When Solution A is mixed with Solution B, the temperature of the mixture increases and a yellow precipitate is observed. What can be concluded from these observations?
- The reaction is thermodynamically favored (spontaneous) at all temperatures.
 - The reaction is thermodynamically favored (spontaneous) only at high temperatures.
 - The reaction is thermodynamically favored (spontaneous) only at low temperatures.
 - The reaction is thermodynamically not favored (spontaneous) at any temperature.

Questions 20-23 refer to the following information:

The removal of copper from metallic ores has been a challenge since ancient times. Strong heating of copper (II) carbonate results in the formation of copper (II) oxide.

Although the decomposition of copper (II) oxide into copper metal and oxygen gas is not thermodynamically favorable under standard conditions, it may be made favorable by adding carbon to the mixture

The resulting reaction is: $2\text{CuO}(s) + \text{C}(s) \rightarrow 2\text{Cu}(s) + \text{CO}_2(g)$



20. Calculate ΔG° for the overall reaction:
- 138kJ/mol
 - 266 kJ/mol
 - 522 kJ/mol
 - 650 kJ/mol
21. Predict and justify the sign of ΔS° for the reaction $\text{CuO}(s) \rightarrow \text{Cu}(s) + \frac{1}{2}\text{O}_2(g)$.
- ΔS is positive because ΔG is positive.
 - ΔS is negative because ΔG is positive.
 - ΔS is positive because the products have more moles of gas than the reactant.
 - ΔS is negative because the products have more moles of gas than the reactant.
22. Why is the overall process thermodynamically favorable?
- Carbon serves as a catalyst to lower the activation energy of the reaction.
 - The process involves coupling a very thermodynamically favorable reaction with one that is not thermodynamically favored.
 - When two reactions are coupled, free energy is always released.
 - The addition of carbon speeds up the reaction because powdered carbon has a large surface area.
23. Which statement is correct about the value of K for the reaction $2\text{CuO}(s) + \text{C}(s) \rightarrow 2\text{Cu}(s) + \text{CO}_2(g)$ at standard conditions?
- K will be less than 1.
 - K will be 0.
 - K will be between 0 and 1.
 - K will be greater than 1.
24. Under normal conditions, an iron nail rusts so slowly that the reaction is not easily observed. What must be true?
- The reaction occurs, but very slowly.
 - The product of the reaction is an invisible gas.
 - The reaction does not occur without a catalyst.
 - The reaction is not thermodynamically favorable.
25. Calculate the ΔG° for the following equation: $2\text{SO}_2(g) + \text{O}_2(g) \rightarrow 2\text{SO}_3(g)$.
- $$\Delta G_f^\circ [\text{SO}_2] = -300 \text{ kJ/mol} \quad \Delta G_f^\circ [\text{SO}_3] = -371 \text{ kJ/mol}$$
- 71kJ/mol
 - 142 kJ/mol
 - 71 kJ/mol
 - 671 kJ/mol
26. How many joules are equivalent to 37.7 cal?
- 9.01J
 - 4.184J
 - 1.51J
 - 158J
27. The quantity of heat that is needed to raise the temperature of a sample of a substance 1.00 degree is called its
- heat capacity
 - specific heat
 - enthalpy
 - kinetic energy
28. Equal masses of two substances, A & B, each absorb 25 Joules of energy. If the temperature of A increases by 4 degrees and the temperature of B increases by 8 degrees, one can say that
- the specific heat of A is double that of B.
 - the specific heat of B is double that of A.
 - the specific heat of B is negative.
 - the specific heat of B is triple that of A.
29. If 25 J are required to change the temperature of 5.0 g of substance A by 2.0°C, what is the specific heat of substance A?

- a. 250 J/g°C b. 63 J/g°C c. 10. J/g°C d. 2.5 J/g°C
30. How much energy is required to change the temperature of 2.00 g aluminum from 20.0°C to 25.0°C? The specific heat of aluminum is 0.902 J/g°C.
- a. 2.3 J b. 9.0 J c. 0.36 J d. 0.090 J
31. Consider the thermal energy transfer during a chemical process. When heat is transferred to the system, the process is said to be _____ and the sign of ΔH is _____.
- a. exothermic, positive c. exothermic, negative
b. endothermic, negative d. endothermic, positive
32. What is the ΔE for a system which has the following two steps:
Step 1: The system absorbs 60J of heat while 40J of work are performed on it.
Step 2: The system releases 30J of heat while doing 70J of work.
- a. 100 J b. 90 J c. 30 J d. zero
33. When two solutions react the container "feels hot." Thus,
- a. the reaction is endothermic.
b. the reaction is exothermic.
c. the energy of the universe is increased.
d. the energy of both the system and the surroundings is decreased.
34. The equation for the standard enthalpy of formation of N_2O_3 is
- a. $N_2O(g) + O_2(g) \rightarrow N_2O_3(g)$ c. $NO(g) + NO_2(g) \rightarrow N_2O_3(g)$
b. $N_2O_5(g) \rightarrow N_2O_3(g) + O_2(g)$ d. $N_2(g) + \frac{3}{2} O_2(g) \rightarrow N_2O_3(g)$
35. For the general reaction, $2 A + B_2 \rightarrow 2 AB$, ΔH is +50.0 kJ. We can conclude that
- a. the reaction is endothermic.
b. the surroundings absorb energy.
c. the standard enthalpy of formation of AB is -50.0 kJ.
d. the molecule AB contains less energy than A or B_2 .
36. Calculate the enthalpy of combustion of C_3H_6 : $C_3H_6(g) + \frac{9}{2}O_2(g) \rightarrow 3CO_2(g) + 3H_2O(l)$ using the following data:
- $3C(s) + 3H_2(g) \rightarrow C_3H_6(g)$ $\Delta H^\circ = 53.3$ kJ
 $C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H^\circ = -394$ kJ
 $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$ $\Delta H^\circ = -286$ kJ
- a. -1517 kJ b. 1304 kJ c. -626 kJ d. -2093 kJ
37. Which one of the following would have an enthalpy of formation value (ΔH_f°) of zero?
- a. $H_2O(g)$ b. $O(g)$ c. $H_2O(l)$ d. $O_2(g)$
38. Calculate the heat of vaporization of titanium (IV) chloride: $TiCl_4(l) \rightarrow TiCl_4(g)$ using the following enthalpies of reaction:
- $Ti(s) + 2Cl_2(g) \rightarrow TiCl_4(l)$ $\Delta H^\circ = -804.2$ kJ
 $TiCl_4(g) \rightarrow 2Cl_2(g) + Ti(s)$ $\Delta H^\circ = 763.2$ kJ
- a. -1567 kJ b. -783.7 kJ c. 1165 kJ d. 41 kJ
39. Calculate the enthalpy of reaction for: $D + F \rightarrow G + M$ using the following equations and data:
- $G + C \rightarrow A + B$ $\Delta H^\circ = +277$ kJ
 $C + F \rightarrow A$ $\Delta H^\circ = +303$ kJ
 $D \rightarrow B + M$ $\Delta H^\circ = -158$ kJ
- a. -132 kJ b. -422 kJ c. +422 kJ d. +132 kJ
40. Calculate the standard enthalpy of the reaction for the process $3NO(g) \rightarrow N_2O(g) + NO_2(g)$ using the standard enthalpies of formation (in kJ/mol): $NO = 90$; $N_2O = 82.1$; $NO_2 = 34.0$
- a. -153.9 kJ b. 206 kJ c. -26.1 kJ d. 386 kJ
41. The standard molar enthalpy of combustion is -1277.3 kJ for the combustion of ethanol.
- $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$
- Calculate the standard molar enthalpy of formation for ethanol based on the following standard enthalpies of formation:
- $\Delta H_f^\circ CO_2 = -393.5$ kJ/mol
 $\Delta H_f^\circ H_2O = -241.8$ kJ/mol
- a. -642.7 kJ/mol b. -235.1 kJ/mol c. 235.1 kJ/mol d. 642.7 kJ/mol
42. Calculate the amount of heat needed to change 25.0g ice at 0°C to water at 0°C. The heat of fusion of $H_2O = 333$ J/g;
- a. 56.5kJ b. 8.33kJ c. 7.06kJ d. 463kJ

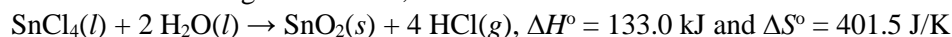
Thermodynamics

43. Which of the following represents an increase in entropy?
- freezing of water
 - boiling of water
 - crystallization of salt from a supersaturated solution
 - the reaction $2\text{NO}(g) \rightarrow \text{N}_2\text{O}_2(g)$
 - the reaction $2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g)$
44. The enthalpy of vaporization of methanol (CH_3OH) is 35.3 kJ/mol at the boiling point of 64.2°C. Calculate the entropy change for methanol going from a liquid to vapor.
600. J/K·mol
 - 105 J/K·mol
 - 551 J/K·mol
 - 551 J/K·mol
 - 105 J/K·mol
45. Calculate the standard entropy change for the following reaction, $\text{Cu}(s) + \frac{1}{2} \text{O}_2(g) \rightarrow \text{CuO}(s)$ given that

$S^\circ [\text{Cu}(s)] = 33.1 \text{ J/K}\cdot\text{mol}$
$S^\circ [\text{O}_2(g)] = 205.14 \text{ J/K}\cdot\text{mol}$
$S^\circ [\text{CuO}(s)] = 42.63 \text{ J/K}\cdot\text{mol}$

- 195.66 J/K
 - 93.09 J/K
 - 45.28 J/K
 - 93.09 J/K
 - 195.66 J/K
46. In which of the following reactions do you expect to have a decrease in entropy?
- $\text{Fe}(s) \rightarrow \text{Fe}(l)$
 - $\text{Fe}(s) + \text{S}(s) \rightarrow \text{FeS}(s)$
 - $2 \text{Fe}(s) + 3/2 \text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s)$
 - $\text{HF}(l) \rightarrow \text{HF}(g)$
 - $2 \text{H}_2\text{O}_2(l) \rightarrow 2 \text{H}_2\text{O}(l) + \text{O}_2(g)$
47. The formation $\frac{1}{2} \text{A}_2 + 2 \text{B}_2 + \text{C} \rightarrow \text{CAB}_4$ has an enthalpy of formation of -104kJ and a change in entropy of -60.8 J/K at 30°C. What is ΔG° and spontaneity of the reaction?
- 85.6kJ, spontaneous
 - 18.3kJ, not spontaneous
 - +18.3kJ, spontaneous
 - +85.6kJ, not spontaneous
 - 85.6kJ, not spontaneous
48. If ΔH° and ΔS are both negative or positive, then ΔG° has a _____ sign.
- positive
 - negative
 - variable
 - large
 - no
49. At what temperature would a given reaction become spontaneous if $\Delta H^\circ = +119 \text{ J}$ and $\Delta S = +263 \text{ J/K}$?
- 452K
 - 2210K
 - 382K
 - 2.21K
 - 363K
50. The free energy change for a given reaction is -36.2 kJ. What is the equilibrium constant at 298K?
- 0.985
 - 2.22×10^6
 - 1.01
 - 8.32×10^{-7}
 - 3.25×10^6

51. Given the following information, calculate ΔG° for the reaction below at 25°C:



- 252.6 kJ
 - 13.4 kJ
 - 13.4 kJ
 - 122.9 kJ
 - 252.6 kJ
51. Given the following information, calculate ΔG° for the reaction below at 25°C: $2 \text{H}_2\text{O}_2(l) \rightarrow 2 \text{H}_2\text{O}(l) + \text{O}_2(g)$

Compound	ΔH° (kJ/mol)	S° (J/K·mol)
$\text{H}_2\text{O}_2(l)$	-187.8	109.6
$\text{H}_2\text{O}(l)$	-285.8	69.9
$\text{O}_2(g)$	—	205.1

- 37700 kJ
 - 342.6 kJ
 - 233.5 kJ
 - 233.5 kJ
 - 157.9 kJ
52. For the process, at 25°C, $\text{I}_2(g) \rightarrow \text{I}_2(s)$, what are the signs of ΔG , ΔH , and ΔS ?
- | | ΔG | ΔH | ΔS |
|----|------------|------------|------------|
| a. | + | - | - |
| b. | - | - | - |
| c. | - | + | + |
| d. | - | - | + |
| e. | + | + | + |
53. If a process is exothermic and not spontaneous, then what must be true?

- a. $\Delta S > 0$ b. $\Delta H > 0$ c. $\Delta G = 0$ d. $\Delta S < 0$ e. $\Delta H = 0$
54. For any reaction at equilibrium, which of the following is true?
 a. $\Delta H < 0$ b. $\Delta S = 0$ c. $\Delta S < 0$ d. $\Delta H = 0$ e. $\Delta G = 0$
55. All of the following have $\Delta G^\circ_f = 0$ EXCEPT
 a. $O_2(g)$ b. $Br_2(g)$ c. $H_2(g)$ d. $Ca(s)$ e. $Hg(l)$

56. Ammonium nitrate spontaneously dissolves in water at room temperature and the process causes the solution to become quite cold. Which of the following is TRUE about the dissolution of ammonium nitrate?
 a. The process is exothermic.
 b. Its solubility will be greater in warmer water.
 c. ΔS° for the reaction is negative.
 d. All solutions of ammonium nitrate are supersaturated.
 e. All solutions of ammonium nitrate are cold.

57. Under which set of conditions is a chemical reaction most likely to be spontaneous?

	ΔH	ΔS	T (temperature)
a.	-	-	low
b.	-	-	high
c.	+	+	low
d.	+	-	high

58. For which reaction do you expect ΔS to be negative?

- a. $2C(s) + O_2(g) \rightarrow 2CO(g)$ c. $H_2O(l, 25^\circ C) \rightarrow H_2O(l, 50^\circ C)$
 b. $Br_2(s) \rightarrow Br_2(l)$ d. $Cl_2(g) + 2HI(g) \rightarrow I_2(s) + 2HCl(g)$

59. Which has the greatest absolute entropy?

- a. one mol of $C(s)$ at $25^\circ C$ c. one mol of $C_2H_6(g)$ at $25^\circ C$
 b. one mol of $CH_3Cl(l)$ at $25^\circ C$ d. one mol of $C_6H_6(l)$ at $25^\circ C$

60. For which of these processes would ΔS° be expected to be the most positive?

- a. $O_2(g) + 2H_2(g) \rightarrow 2H_2O(g)$ c. $N_2O_4(g) \rightarrow 2NO_2(g)$
 b. $H_2O(l) \rightarrow H_2O(s)$ d. $NH_4NO_2(s) \rightarrow N_2(g) + 2H_2O(g)$

61. For which of these processes is the value of ΔS expected to be negative?

- I. Sugar is dissolved in water
 II. Steam is condensed
 III. $CaCO_3$ is decomposed into CaO and CO_2 .

- a. I only c. II only
 b. I and III only d. II and III only

62. Which set of conditions is most likely to result in a reaction that is spontaneous as written?

	ΔH	ΔS	T
a.	< 0	< 0	$500^\circ C$
b.	< 0	< 0	$0^\circ C$
c.	> 0	< 0	$0^\circ C$
d.	> 0	< 0	$500^\circ C$

63. For which of these processes is the sign of the enthalpy change different from the others?

- e. $Al_2O_3(s) \rightarrow 2Al(s) + 3/2O_2(g)$ g. $Cl_2(g) \rightarrow 2Cl(g)$
 f. $H_2O(s) \rightarrow H_2O(l)$ h. $Cl(g) + e^- \rightarrow Cl^-(g)$

64. For the process $O_2(g) \rightarrow 2O(g)$, $\Delta H^\circ = +498$ kJ. What would be predicted for the sign of ΔS_{rxn} and the conditions under which this reaction would be spontaneous?

	ΔS_{rxn}	Spontaneous
a.	positive	at low temperatures only
b.	positive	at high temperatures only
c.	negative	at high temperatures only
d.	negative	at low temperatures only

65. For the reaction $NH_4Cl(s) \rightarrow NH_3(g) + HCl(g)$, $\Delta H^\circ = +176$ kJ and $\Delta G^\circ = +91.2$ kJ at 298 K. What is the value of ΔG at 1000 K?

- i. -109 kJ j. -64 kJ k. +64 kJ l. +109 kJ

66. When ammonium nitrate, $NH_4NO_3(s)$ is added to water at $25^\circ C$, it dissolves spontaneously and the temperature of the solution decreases. This indicates that the factor causing the substance to dissolve is a change in

- m. energy n. enthalpy o. entropy p. temperature

