

AP Chem Unit 11 Thermodynamics

Practice Exam - KEY

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- 1 C In an exothermic reaction, energy is given off as the products are created because the products have less potential energy than the reactants.
- 2 C The bond energy is the energy that must be put into a bond to break it. First let's figure out how much energy must be put into the reactants to break their bonds.
To break 2 moles of H-H bonds, it takes $(2)(500) \text{ kJ} = 1,000 \text{ kJ}$
To break 1 mole of O=O bonds, it takes 500 kJ.
So to break up the reactants, it takes +1,500 kJ.
Energy is given off when a bond is formed; that's the negative of the bond energy. Now let's see how much energy is given off when 2 moles of H₂O are formed.
2 moles of H₂O molecules contain 4 moles of O-H bonds, so $(4)(-500) \text{ kJ} = -2,000 \text{ kJ}$ are given off.
So the value of ΔH° for the reaction is
 $(-2,000 \text{ kJ, the energy given off}) + (1,500 \text{ kJ, the energy put in}) = -500 \text{ kJ}$.
- 3 A Remember, $\Delta G = \Delta H - T \Delta S$. If the reaction is spontaneous at standard temperature but becomes nonspontaneous at higher temperatures, then ΔG is negative only at lower temperatures. This can only happen when ΔH is negative (which favors spontaneity) and ΔS is negative (which favors nonspontaneity). As the value of T increases, the influence of ΔS increases, eventually making the reaction nonspontaneous.
- 4 C Point 3 represents the activated complex, which is the point of highest energy. This point is the transition state between the reactants and the products.
- 5 D The reaction is spontaneous, so ΔG must be less than zero, so (I) is not true. The reaction is endothermic, so ΔH must be greater than zero; therefore, (II) is true.
The only way that an endothermic reaction can be spontaneous is if the entropy is increasing, so ΔS is greater than zero, and (III) is true.
- 6 C The addition of a catalyst speeds up a reaction by lowering the activation energy. A catalyst has no effect on the entropy or enthalpy change of a reaction.
- 7 D The equations given above the question give the heats of formation of all the reactants and products (remember, the heat of formation of O₂, an element in its most stable form, is zero).
 ΔH° for a reaction = $(\Delta H^\circ$ for the products) - $(\Delta H^\circ$ for the reactants).
First, the products:
From 2 H₂O, we get $2z$
From CO₂, we get y
So ΔH° for the products = $2z + y$
Now the reactants:
From CH₄, we get x . The heat of formation of O₂ is defined to be zero, so that's it for the reactants.
 ΔH° for the reaction = $(2z + y) - (x) = 2z + y - x$.

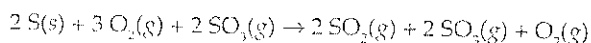
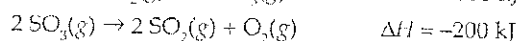
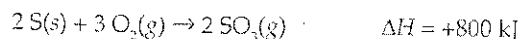
Practice Exam - Key

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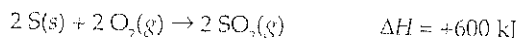
8 C The reaction that forms 2 moles of HF(g) from its constituent elements has an enthalpy change of -540 kJ. The heat of formation is given by the reaction that forms 1 mole from these elements, so you can just divide -540 kJ by 2 to get -270 kJ.

9 E Ice melts spontaneously at room temperature, so ΔG is less than zero; therefore, (I) is not true. In this reaction, ice melts, so heat is absorbed to break up the intermolecular forces, and ΔH is greater than zero; therefore, (II) is true. Liquid water is less orderly than ice, so the entropy change when ice melts is positive; therefore, ΔS is greater than zero, and (III) is true.

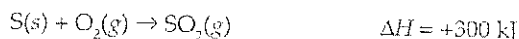
10 A You can use Hess's law. Add the two reactions together, and cancel things that appear on both sides.



This reduces to



Now we can cut everything in half to get the equation we want.



11. ANS: D

Realize this is a combustion reaction and heat energy is lost. Vaporization is an endothermic process and *requires* energy. That means some of the energy lost by the combustion system is subsequently reused by the system to vaporize water if $\text{H}_2\text{O}(g)$ is formed. This lowers the total energy yield.

As a result, more total energy is lost from the combustion reaction that produces *liquid* water compared to gaseous water. So, expect that this reaction would yield $2(-44 \text{ kJ mol}^{-1})$ more energy if liquid water is formed rather than water vapor. Both answers C and D yield more energy, but C didn't factor in that two moles of water are involved.

DIF: Hard OBJ: 5.6 TOP: Thermochemistry

MSC: 1999 #61 NOT: 31% answered correctly

12 ANS: D

A negative ΔS value implies a decrease in disorder, therefore a more ordered product or set of products is/are formed. Look for reactions that produce more condensed states and/or fewer moles of product.

Reaction 1: 2 moles solid \rightarrow 4 moles solid...therefore, more disorder... + 3 moles gas...again, more disorder

Reaction 2: 3 moles of ions \rightarrow 1 mole of solid...far more ordered

Reaction 3: 2 moles of gas \rightarrow 1 mole of gas...more ordered even though all are gases

DIF: Medium OBJ: 5.12 TOP: Thermodynamics

MSC: 1994 Q#35 NOT: 54% answered correctly

1. (a) Use the entropy values in the table.

$$\Delta S^\circ = \sum S^\circ_{\text{products}} - \sum S^\circ_{\text{reactants}}$$

$$\Delta S^\circ = [(6)(213.6) + (6)(69.9)] - [(212.13) + (6)(205)] \text{ J/K}$$

$$\Delta S^\circ = 259 \text{ J/K}$$

- (b) Use the equation below. Remember that enthalpy values are given in kJ and entropy values are given in J.

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

$$\Delta G^\circ = (-2,801 \text{ kJ}) - (298)(0.259 \text{ kJ}) = -2,880 \text{ kJ}$$

- (c) Use the equation below. Remember that the gas constant is given in terms of J.

$$\log K = \frac{\Delta G^\circ}{-2.303RT}$$

$$\log K = \frac{(-2,880,000)}{(-2.303)(8.31)(298)} = 505$$

$$K = 10^{505}$$

- (d) The enthalpy change of the reaction, H° , is a measure of the energy given off by 1 mole of glucose.

$$\text{Moles} = \frac{\text{grams}}{\text{MW}}$$

$$\text{Moles of glucose} = \frac{(1.00 \text{ g})}{(180 \text{ g/mol})} = 0.00556 \text{ moles}$$

$$(0.00556 \text{ mol})(2,801 \text{ kJ/mol}) = 15.6 \text{ kJ}$$

- 2 (a) ΔS° is negative because the products are less random than the reactants. That's because gas is converted into liquid in the reaction.

(b) The value of ΔS° would increase, becoming less negative because $\text{H}_2\text{O}(g)$ is more random than water but remaining negative because the entropy would still decrease from reactants to products.

(c) ΔG° is negative because the reaction proceeds spontaneously.

(d) ΔH° must be negative at 298 K. For a reaction to occur spontaneously from standard conditions, either ΔS° must be positive or ΔH° must be negative. This reaction is spontaneous although ΔS° is negative, so ΔH° must be negative.