|  |  |
| --- | --- |
| **Name:** |  |
| **Your Score:** |  |
| **Maximum Score:** | **40** |

**AP CHEMISTRY**

**UNIT 06ABCDEFGHI & 09ABCDEFG TEST**

## Instructions

* **Write your name at the top of this page**
* **You are provided with an equations & constants sheet, and a periodic table**
* **In SECTION A you will find multiple-choice questions. Answer these by choosing the one letter that corresponds to the best answer**
* **In SECTION B you will find a number of free response questions. Write your answers in the spaces provided. The number of points for each part of each question is shown in parentheses**
* **Attempt ALL the questions**
* **Where appropriate show ALL working and pay attention to units and significant figures**
* **Present all your work as neatly as possible**

**SECTION A: Multiple-Choice (Select the best answer)**

????1. An endothermic chemical reaction has which combination of conditions? (6.1 & 9.2)

(A) Negative ΔH, an increase of the temperature of the surroundings, and positive ΔS

(B) Positive ΔH, an increase of the temperature of the surroundings, and negative ΔS

(C) Positive ΔH, a decrease of the temperature of the surroundings, and either a positive or a negative ΔS

(D) Positive ΔH, an increase of the temperature of the surroundings, and either a positive or a negative ΔS

Questions 2 through 4

??2. Copper metal has a specific heat of 0.3850 J g-1 °C -1. Calculate the amount of energy required to raise the temperature of 23.00 g of Cu metal from 20.00 °C to 75.00 °C. (6.4)

(A) 487.0 J

(B) 487.0 kJ

(C) 664.1 J

(D) 664.1 kJ

3. When the hot copper metal at 75.00 °C described in question 2 is lowered into a calorimeter filled with cold water at an initial temperature of 25 °C, what will the new, final temperature in the calorimeter be? (Assume no heat loss, and that the specific heat capacity of water is 4.184 J g-1 oC-1). (6.4)

(A) Exactly halfway between 25 °C and 75 °C

(B) Closer to 25 °C than 75 °C

(C) Closer to 75 °C than 25 °C

(D) 100 °C

4. The process of the contents of the calorimeter reaching a final, single temperature is best desribed as (6.3)

(A) Enthalpy of vaporization

(B) Thermal equilibrium

(C) Specific heat capacity

(D) Enthalpy of fusion

5. The combustion of liquid ethanol (shown below) has a ΔH° = - 278 kJ/molrxn

C2H5OH(l) + 3O2(g) 🡺 2CO2(g) + 3H2O(l)

what is ΔH° for the reaction shown below? (6.9)

4CO2(g) + 6H2O(l) 🡺 2C2H5OH(l) + 6O2(g)

(A) + 556 kJ/molrxn

(B) + 278 kJ/molrxn

(C) - 556 kJ/molrxn

(D) + 1.29 x 10-5 kJ/molrxn

??6. Arrange the following compounds in order of **increasing** absolute entropies. (9.2)

C3H8(g), CH4(g), NaCl(s), and C2H5OH(l)

(A) NaCl(s) < C2H5OH(l) < C3H8(g) < CH4(g)

(B) CH4(g) < C2H5OH(l) < C3H8(g) < NaCl(s)

(C) NaCl(s) < C3H8(g) < CH4(g) < C2H5OH(l)

(D) NaCl(s) < C2H5OH(l) < CH4(g) < C3H8(g)

7. The ΔS° when a solid is converted to a gas is (9.2)

(A) always negative

(B) always positive

(C) sometimes is positive and sometimes is negative it depends on the temperature

(D) it is not possible to tell from the data given

8. A positive sign for ΔG° for a reaction indicates that (9.3 & 9.4)

(A) the reaction is endothermic

(B) ΔS° must be < 0

(C) the reaction is slow

(D) the reaction is thermodynamically unfavored

??9. Which of the following has an enthalpy of formation ΔHf° = 0? (6.8)

(A) Na

(B) O

(C) N

(D) H

10. Reactions that have positive ΔH° and negative ΔS° are (9.3)

(A) never thermodynamically favored

(B) always thermodynamically favored

(C) thermodynamically favored at high temperatures

(D) thermodynamically favored at low temperatures

**ANSWERS**

**1. C**

**2. A**

**3. B**

**4. B**

**5. A**

**6. D**

**7. B**

**8. D**

**9. A**

**10. A**

**SECTION B Free Response Questions (Write answers in the spaces provided)**

# Question 1 (2011, 3 (a)-(c))

Hydrogen gas burns in air according to the equation below.

2H2(g) + O2(g) 🡺 2H2O(l)

(a) Calculate ΔHo, for the reaction represented by the equation above. (2) (6.6 & 6.8)

(The molar enthalpy of formation, ΔHof , for H2O(l) is − 285.8 kJ mol−1)

The equation represents the enthalpy of formation of 2 moles of H2O(l) so, (2)(-285.8) -571.6 kJ mol-1

(b) Calculate the amount of energy in kJ, that is released when 10.0 g of H2(g) is burned in air. (2) (6.6)

10.0 g/2.016 g mol-1 = moles of H2

(10.0 g/2.016 g mol-1)(285.8 kJ/mol) = 1.42 x 103 kJ released

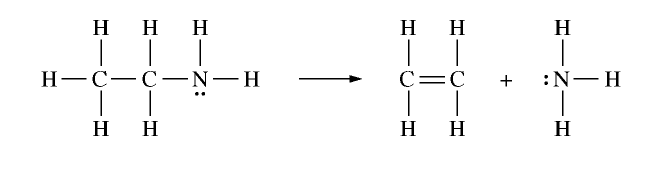
(c) Given that the molar enthalpy of vaporization, ΔHovap, for H2O(l) is + 44.0 kJ mol−1 at 298 K, what is the standard enthalpy change, ΔHo298, for the reaction, 2H2(g) + O2(g) 🡺 2H2O(g)? (3) (6.9)

2H2(g) + O2(g) → 2H2O(l) −571.6 kJ  
2H2O(l) → 2H2O(g)  +2(44.0) kJ

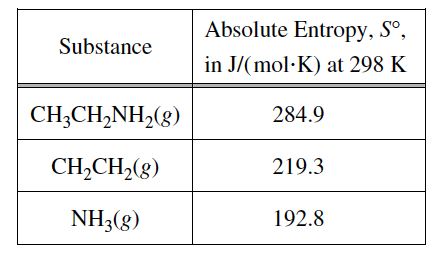
Sum these equations to yield,

2H2(g) + O2(g) → 2H2O(g) −483.6 kJ

**Question 2 (2012, 3(a)-(c))**



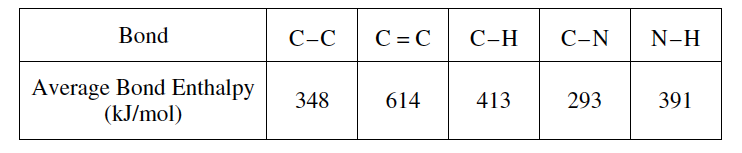
A sample of CH3CH2NH2 is placed in an insulated container, where it decomposes into ethene and ammonia according to the reaction represented above.



(a) Using the data in the table above, calculate the value, in J/(molrxn K), of the standard entropy change ΔS°, for the reaction at 298 K. (1) (9.2)

ΔS° = (219.3 + 192.8) – 284.9 = 127.2 J/(molrxn K)

(b) Using the data in the table below, calculate the value, in kJ/molrxn, of the standard enthalpy change, ΔH°, for the reaction at 298 K. (2) (6.7)



Bonds broken (endothermic, POSITIVE)

= +[5(413) + 293 + 348 + 2(391)] = +3488

Bonds made (Exothermic, NEGATIVE)

= -[4(413) + 614 + 3(391)] = -3439

Sum of Endo and Exo processes = ΔH° = 49 kJ/molrxn

(c) Based on your answer to part (b), predict whether the temperature of the contents of the insulated container will increase, decrease, or remain the same as the reaction proceeds. **Justify your prediction**. (2) (6.1)

Decrease, since an endothermic process absorbs energy from the surroundings.

**Question 3 (Based on 2006B, 3 (a), (d) and (f))**

Answer the following questions about the thermodynamics of the reactions represented below.

**Reaction *X*: ½ I2(*s*) + ½C*l*2(*g*) 🡺 IC*l*(*g*)** Δ*Hf°*298= 18 kJ mol-1, Δ*S°*298 = 78 J K-1 mol-1

**Reaction *Y*: ½ I2(*s*) + ½ Br2(*l*) 🡺 IBr(*g*)** Δ*Hf°*298 = 41 kJ mol-1, Δ*S°*298 = 124 J K-1 mol-1

(a) Is reaction *X,* represented above, thermodynamically favored under standard conditions? Justify your answer with a calculation. (2) (9.3)

Yes, since it has a negative ∆G°

∆G° = ∆H° − T∆S° = (18 kJ/mol) − ( 298 K)(0.078 kJ/molK) = −5 kJ/mol

(b) Explain why the standard entropy change is greater for reaction *Y* than for reaction *X*. (2) (9.1 & 9.2)

The conversion of a liquid into a gas (Y) is a more profound change in dispersal than a gas into a gas (X)

(c) For the vaporization of solid iodine, I2(*s*) 🡺 I2(*g*), the value of Δ*H*θ298 is 62 kJ mol-1. Using this information, calculate the value of Δ*H*θ298 for the reaction represented below. (2) (6.9)

I2(*g*) + Cl2(*g*) 🡺 2IC*l* (*g*)

I2(s) + Cl2(g) → 2ICl(g) (2)(18) kJ  
I2(g) → I2(s) −62 kJ

Sum these equations to yield,

I2(g) + Cl2(g) → 2ICl(g) − 26 kJ

**Question 4**

(a) Calculate the ΔH°rxn for the decomposition reaction below given the data that follow. (3) (6.9)

2H3BO3(aq) → B2O3(s) + 3H2O(l)

1. H3BO3(aq) → HBO2(aq) + H2O(l) ΔH° = - 0.0200 kJ/molrxn
2. H2B4O7(aq) + H2O(l) → 4HBO2(aq) ΔH° = - 11.3 kJ/molrxn
3. H2B4O7(aq) → 2B2O3(s) + H2O(l) ΔH° = + 17.5 kJ/molrxn

Double #1, half and reverse #2, half #3, then add.

ΔH°rxn = 2(-0.0200) + (+11.3/2) + (+17.5/2) = +14.4 kJ/molrxn

(b) Assuming that the decomposition reaction is thermodynamically favorable, does your answer in (a) suggest that the decomposition reaction is driven *at least in part,* by enthalpy? Explain your answer. (2) (9.3)

No. A positive enthalpy does not help to make the ΔG° for the reaction negative and therefore

thermodynamically favored.

(c) If the decomposition reaction were to create gaseous H2O instead of liquid H2O, would the entropy change be more positive or more negative. Explain your answer. (2) (9.1 & 9.2)

More positive.

ΔS° = S°products – S°reactants, with a more positive value for S°products that gaseous H2O would bring, the ΔS°

will be more positive.

**Question 5**

Consider the individual transformations shown below, that represent the changes that take place

when solid sodium bromide is formed from its elements. (The variables S through X represent the

numerical values for the enthalpy changes in each case and may be positive OR negative).

|  |  |
| --- | --- |
| **Transformation** | **ΔH° in kJ mol-1** |
| Na(s) 🡺 Na(g) | S |
| Na(g) 🡺 Na+(g) + e- | T |
| Na+(g) 🡺 Na2+(g) + e- | U |
| Br2(l) 🡺 2Br(g) | V |
| Br(g) + e- 🡺 Br -(g) | W |
| Na+(g) + Br -(g) 🡺 NaBr(s) | X |

The ΔHof for NaBr(s) = -361.0 kJ mol-1

(a) Given that the formation of NaBr(s) from its elements is a thermodynamically favored process, is it enthalpy or entropy that drives the formation of NaBr(s) from its elements? Explain. (2) (9.3)

Enthalpy.

A negative enthalpy helps to make the ΔG° for the reaction negative and therefore thermodynamically

favored. The reaction probably has a negative ΔS° (a solid and a liquid becoming a solid), which would

not help to favor a negative ΔG°.

(b) Express ΔHof for NaBr(s) as an algebraic expression made up of the letters S through X. (1) (6.9)

ΔHof for NaBr(s) = S + T + V/2 + W + X

(c) Name the process illustrated in the chemical equation with the ΔH° value = X. (1) (3.3 as it relates to 6.8)

Lattice energy/enthalpy

(d) Four of the transformations in S through X are endothermic, and two are exothermic. Which four of the processes labeled S through X, would you expect to be endothermic? (1) (6.1)

S, T, U and V