**Name: Period: Seat#:**

**Worksheet #6**

**Directions:** Show all work in a way that would earn you credit on the AP Test! This is always the rule! Grading rubrics posted in the Google Answer Key Drive. Check your work, correct in green pen after you try them yourself in an honest way! Don’t peek at rubrics while you work! **USE BINDER PAPER, STAPLE TO YOUR WORKSHEET**. Clearly label work.

**LONG ASSIGNMENT! DON’T WAIT UNTIL THE LAST MINUTE! BREAK IT INTO CHUNKS!**

**SET A TIMER FOR 1.5 MIN PER FRQ PART AND SEE IF YOU FINISH ON TIME!**

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| 2007B | 1. A sample of solid U3O8 is placed in a rigid 1.500 L flask. Chlorine gas, Cl2*(g)*, is added and the flask is heated to 862°C. The equation for the reaction that takes place and the equilibrium constant expression for the reaction are given below. U3O8*(s)* + 3 Cl2*(g)* ↔ 3 UO2Cl2*(g)* + O2*(g)* $K\_{p}=\frac{(P\_{UO\_{2}Cl\_{2}})^{3}(P\_{O\_{2}})}{(P\_{Cl\_{2}})^{3}}$When the system is at equilibrium the partial pressure of the Cl2*(g)* is 1.007 atm and the partial pressure of UO2Cl2*(g)* is 9.734 x 10-4 atm. 1. Calculate the partial pressure of O2*(g)* at equilibrium at 862°C.
2. Calculate the value of the equilibrium constant, Kp, for the system at 862°C.
3. Calculate the Gibbs free-energy change, ∆G°, for the reaction at 862°C.
4. State whether the entropy change, ∆S°, for the reaction at 862°C is positive, negative, or zero. Justify answer.
5. State whether the enthalpy change, ∆H°, for the reaction at 862°C is positive, negative, or zero. Justify answer.
6. After a certain period of time, 1.00 mol of O2*(g)* is added to the mixture in the flask. Does the mass of U3O8*(s)* in the flask increase, decrease, or remain the same? Justify your answer.
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| 2007 | N2*(g)* + 3 F2*(g)* 🡪 2 NF3*(g)* ∆H°298 = – 264 kJ mol-1 ; ∆S°298 = – 278 J K-1 mol-12. The following questions relate to the synthesis reaction represented by the equation in the box above. (a) Calculate the value of the standard free energy change, ∆G°298 , for the reaction. (b) Determine the temperature at which the equilibrium constant, Keq, for the reaction is equal to 1.00. (Assume that ∆H° and ∆S° are independent of temperature.)(c) Calculate the standard enthalpy change, ∆H°, that occurs when a 0.256 mol sample of NF3*(g)* is formed from N2*(g)* and F2*(g)* at 1.00 atm and 298 K. The enthalpy change in a chemical reaction is the difference between energy absorbed in breaking bonds in the reactants and energy released by bond formation in the products. (d) How many bonds are formed when two molecules of NF3 are produced according to the equation in the box above?(e) Use both the information in the box above and the table of average bond enthalpies below to calculate the average enthalpy of the F – F bond.

|  |  |
| --- | --- |
| Bond | Average Bond Enthalpy(kJ mol-1) |
| N ≡ N | 946 |
| N – F | 272 |
| F – F | ? |

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| 2006B | 3. Answer the following questions about the thermodynamics of the reactions represented below. Reaction X: $\frac{1}{2}I\_{2}\left(s\right)+\frac{1}{2}Cl\_{2}\left(g\right)\leftrightarrow ICl\left(g\right) ∆H\_{f}^{°}=18 kJ mol^{-1}, ∆S\_{298}^{°}=78 J K^{-1}mol^{-1}$Reaction Y: $\frac{1}{2}I\_{2}\left(s\right)+\frac{1}{2}Br\_{2}\left(l\right)\leftrightarrow IBr\left(g\right) ∆H\_{f}^{°}=41 kJ mol^{-1}, ∆S\_{298}^{°}=124 J K^{-1}mol^{-1}$1. Is reaction X, represented above, spontaneous under standard conditions? Justify your answer with a calculation.

 1. Calculate the value of the equilibrium constant, Keq, for reaction *X* at 25°C
2. What effect will an increase in temperature have on the equilibrium constant for reaction *X*? Explain your answer.
3. Explain why the stnadard entropy change is greater for reaction *Y* than for reaction *X*.
4. Above what temperature will the value of the equilibrium constant for reaction *Y* be greater than 1.0? Justify your answer with calculations.
5. For the vaporization of solid iodine, I2(s) 🡪 I2(g), the value of ∆H°298 is 62 kJ mol-1. Using this information, calculate the value of ∆H°298 for the reaction represented below.

I2(g) + Cl2(g) ↔ 2 ICl(g) |
| 2006 |  $CO\left(g\right)+\frac{1}{2}O\_{2}\left(g\right)\leftrightarrow CO\_{2}\left(g\right) $2. The combustion of catrbon monoxide is represented by the equation above. 1. Determine the value of the standard enthalpy change $∆H\_{rxn}^{°}$, for the combustion of CO*(g)* at 298 K using the following information. $$C\left(s\right)+\frac{1}{2}O\_{2}\left(g\right)\leftrightarrow CO\left(g\right) ∆H\_{298}^{°}=-110.5 kJ mol^{-1}$$

$$C\left(s\right)+O\_{2}\left(g\right)\leftrightarrow CO\_{2}\left(g\right) ∆H\_{298}^{°}=-393.5 kJ mol^{-1}$$1. Determine the value of the standard entropy change, $∆S\_{rxn}^{°}$, for the combustion of CO*(g)* at 298 K using the information in the following table.

|  |  |
| --- | --- |
| Substance | $∆S\_{rxn}^{°}$ (J mol-1K-1) |
| CO*(g)* | 197.7 |
| CO2*(g)* | 213.7 |
| O2*(g)* | 205.1 |

1. Determine the standard free energy change, $∆G\_{rxn}^{°} $, for the reaction at 298 K. Include units with your answer.
2. Is the reaction spontaneous under standard conditions at 298 K ? Justify your answer.
3. Calculate the value of the equilibrium constant, K*eq*, for the reaction at 298 K.
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| 2004B | $N\_{2}\left(g\right)+2 H\_{2}\left(g\right)\leftrightarrow N\_{2}H\_{4}\left(g\right) ∆H\_{298}^{°}=+95.4 kJ mol^{-1}$ ∆S°298 = – 176 J K-1 mol-17. Answer the following questions about the reaction represented above using principles of thermodynamics. 1. On the basis of the thermodynamic data given above, compare the sum of the bond strengths of the reactants to the sum of the bond strengths of the product. Justify your answer.
2. Does the entropy change of the reaction favor the reactants or the product? Justify your answer.
3. For the rxn under the conditions specified, which is favored, the reactant or the product? Justify your answer.
4. Explain how to determine the value of the equilibrium constant, K*eq* , for the reaction. (Do not do any calculations.)

(e) Predict whether the value of K*eq* for the reaction is greater than 1, equal to 1, or less than 1. Justify your answer.  |
| 2004 |  $2 Fe\left(s\right)+\frac{3}{2}O\_{2}\left(g\right)\rightarrow Fe\_{2}O\_{3}\left(s\right) ∆H\_{f}^{°}=-824 kJ mol^{-1}$2. Iron reacts with oxygen to produce iron(III) oxide, as represented by the equation above. A 75.0 g sample of Fe*(s)* is mixed with 11.5 L of O2*(g)* at 2.66 atm and 298 K. 1. Calculate the number of moles of each of the following before the reaction begins.
	1. Fe*(s)*
	2. O2*(g)*
2. Identify the limiting reactant when the mixture is heated to produce Fe2O3*(g)*. Support your answer with calculations.
3. Calculate the number of moles of Fe2O3*(s)* produced when the reaction proceeds to completion.
4. The standand free energy of formation, $∆G\_{f}^{°}$ , of Fe2O3*(s)* is $-740. kJ mol^{-1}$ at 298 K.
	1. Calculate the standard entropy of formation, $∆S\_{f}^{°}$ , of Fe2O3*(s)* at 298 K. Include units with your answer.
	2. Which is more responsible for the spontaneity of the formation reaction at 298 K, the standard enthalpy of formation, $∆H\_{f}^{°}$ , or the standard entropy of formation, $∆S\_{f}^{°}$ ? Justify your answer.

The reaction represented below also produces iron(III) oxide. The value of $∆H^{°}$ for the reaction is $-280. $kJ per mole of Fe2O3*(s)* formed.  $2 FeO\left(s\right)+\frac{1}{2}O\_{2}\left(g\right)\rightarrow Fe\_{2}O\_{3}\left(s\right)$1. Calculate the standard enthalpy of formation, $∆H\_{f}^{°}$ , of FeO*(s).*
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| 2003 | 7. Answer the following questions that relate to the chemistry of nitrogen. 1. Two nitrogen atoms combine to form a nitrogen molecule, as represented by the following equation.

2 N*(g)* 🡪 N2*(g)*Using the table of average bond energies below, determine the enthalpy change, ∆*H* , for the reaction.

|  |  |
| --- | --- |
| Bond | Average Bond Enthalpy(kJ mol-1) |
| N – N | 160 |
| N = N | 420 |
| N ≡ N | 950 |

1. The reaction between nitrogen and hydrogen to form ammonia is represented below.

$$N\_{2}\left(g\right)+3 H\_{2}\left(g\right)\rightarrow 2 NH\_{3}\left(g\right) ∆H\_{}^{°}=-92.2 kJ$$Predict the sign of the standard entropy change, ∆*S*° , for the reaction. Justify your answer. 1. The value of ∆*G*° for the reaction represented in part (b) is negative at low temperatures but positive at high temperatures. Explain.
2. When N2*(g)* and H2*(g)* are placed in a sealed container at a low temperature, no measurable amoung of NH3*(g)* is produced. Explain.
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| 2002B*2002B cont.* | 3. Nitrogen monoxide, NO*(g)*, and carbon monoxide, CO*(g),* are air pollutants generated by automobiles. It has been proposed that under suitable conditions these two gases could react to form N2*(g)* and CO2*(g)*, which are components of unpolluted air. 1. Write a balanced equation for the reaction described above. Indicate whether the carbon in CO is oxidized or whether it is reduced in the reaction. Justify your answer.
2. Write the expression for the equiibrium constant, *Kp*, for the reaction.
3. Consider the following therodynamic data.

 NO CO CO2∆$G\_{f}^{°}$ (kJ mol-1) +86.55 -137.15 -394.361. Calculate the value of ∆*G*° for the reaction at 298 K.
2. (ii) Given that the ∆*H*° for the reaction at 298 K is -746 kJ per mole N2*(g)* formed, calculate the value of ∆*S*° for the reaction at 298 K. Include units with your answer.
3. For the reaction at 298 K, the value of *Kp* is 3.3 x 10120. In an urban area, typical pressures of the gases in the reaction are PNO = 5.0 x 10-7 atm, PCO = 5.0 x 10-5 atm, PN2 = 0.781 atm, and PCO2 = 3.1 x 10-4 atm.
	1. Calculate the value of ∆*G* for the rxn at 298 K when the gases are at the partial pressures given above.
	2. In which direction (to the right or to the left) will the reaction be spontaneous at 298 K with these partial pressures? Explain.
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| 2001 |  $ 2NO\left(g\right)+O\_{2}\left(g\right)\rightarrow 2 NO\_{2}\left(g\right) ∆H\_{}^{°}=-114.1 kJ, ∆S\_{}^{°}=-146.5 J K^{-1}$2. The reaction represented above is one that contributes significantly to the formation of photochemical smog. 1. Calculate the quantity of heat released when 73.1 g of NO*(g)* is converted to NO2*(g).*
2. For the reaction at 25°C, the value of the standard free-energy change, ∆*G*°, is -70.4 kJ.
	1. Calculate the value of the equilibrium constant, *Keq*, for the reaction at 25°C.
	2. Indicate whether the value of ∆*G*° would become more negative, less negative, or remain unchanged as the temperature is increased. Justify your answer.
3. Use the data in the table below to calculate the value of the standard molar entropy, ∆S° , for O2(g) at 25°C

|  |  |
| --- | --- |
|  | Standard Molar Entropy, *S°*(J K-1 mol-1) |
| NO*(g)* | 210.8 |
| NO2*(g)* | 240.1 |

1. Use the data in the table below to calculate the bond energy, in kJ mol-1 , of the nitrogen-oxygen bond in NO2. Assume that the bonds in the NO2 molecule are equivalent (i.e., they have the same energy).

|  |  |
| --- | --- |
|  | Bond Energy(kJ mol-1) |
| Nitrogen-oxygen bond in NO | 607 |
| Oxygen-oxygen bond in O2 | 495 |
| Nitrogen-oxygen bond in NO2 | ? |

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| 1999 | 6. Answer the following questions in terms of thermodynamic principles and concepts of kinetic molecular theory. 1. Consider the reaction represented below, which is spontaneous at 298 K.

 $ CO\_{2}\left(g\right)+2 NH\_{3}\left(g\right)\rightarrow CO\left(NH\_{2}\right)\_{2}\left(s\right)+H\_{2}O(l) ∆H\_{298}^{°}=-134 kJ$1. For the reaction, indicate whether the standard entropy change ∆$S\_{298}^{°}$ , is positive, or negative, or zero. Justify your answer.
2. Which factor, the change in enthalpy $∆H\_{298}^{°}$ , or the change in entropy $∆S\_{298}^{°}$ , provides the principal driving force for the reaction at 298 K ? Explain.
3. For the reaction, how is the value of the standard free energy change, ∆*G*° , affected by an increase in temperature? Explain.
4. Some reactions that are predicted by their sign of ∆*G*° to be spontaneous at room temperature do not proceed at a measurable rate at room temperature.
	1. Account for this apparent contradiction.
	2. A suitable catalyst increases the rate of such a reaction. What effect does the catalyst have on ∆*G*° for the reaction? Explain.
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| **Reflection:** Think about the types of mistakes you made, things you need to restudy, things that tricked you, etc. One of the most important skills to develop in AP Chem is self reflection and not making the same mistakes. The joke is – you should always make NEW mistakes, not the SAME mistakes ☺ |