Name ______________________________ **AP Chemistry**

Chapter 16 HW 2: Due 3/4/16. Complete all free response questions. One will be graded. Show all work. Box and clearly label all final free response answers. #YOTAPCO

1. 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 $N_2(g)$ and $CO_2(g)$, which are components of unpolluted air.
(a) 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.
(b) Write the expression for the equilibrium constant, K_p, for the reaction.
(c) Consider the following thermodynamic data.

	NO	CO	CO_2
ΔG_f°(kJ mol^{-1})	+86.55	−137.15	−394.36

(i) Calculate the value of ΔG° for the reaction at 298 K.
(ii) Given that ΔH° for the reaction at 298 K is −746 kJ per mole of $N_2(g)$ formed, calculate the value of ΔS° for the reaction at 298 K. Include units with your answer.

(d) For the reaction at 298 K, the value of K_p is 3.3 x 10^{120}. In an urban area, typical pressures of the gases in the reaction are P_{NO} = 5.0 x 10^{-7} atm, P_{CO} = 5.0 x 10^{-5} atm, P_{N2} = 0.781 atm, and P_{CO2} = 3.1 x 10^{-4} atm.
(i) Calculate the value of ΔG for the reaction at 298 K when the gases are at the partial pressures given above.
(ii) In which direction (to the right or to the left) will the reaction be spontaneous at 298 K with these partial pressures? Explain.

1
a. $2NO + 2CO \rightarrow 2CO_2 + N_2$

1
CO is oxidized because it loses e⁻ as Carbon's charge goes from +2 to +4.

1
b. $K_p = \dfrac{(pCO_2)^2(pN_2)}{(pNO)^2(pCO)^2}$

c.i $\Delta G = [2(-394.36)+0] - [2(86.55)+2(-137.15)]$

$\Delta G = -788.72 - [-101.2]$

2
$\boxed{\Delta G = -687.5\text{ KJ}}$

c.ii $\Delta G = \Delta H - T\Delta S$

$-687500 = -746000 - 298\Delta S$

$\boxed{\Delta S = -0.196\text{ KJ/K}}$ 2

d.i $Q = \dfrac{(3.1\times10^{-4})^2(0.781)}{(5.0\times10^{-7})^2(5.0\times10^{-5})^2}$ $Q = 1.20\times10^{14}$

$\Delta G = \Delta G^\circ + RT\ln Q$

~~-688800~~ $\Delta G = -687500 + (8.3145)(298)(\ln 1.20\times10^{14})$

$\Delta G = -687500 + 80324$

2
$\boxed{\Delta G = -607\text{KJ}}$

d.ii
1
ΔG is negative & K_p is > 1 so the reaction is spontaneous in the forward (right) direction.

2. $CO(g) + 2\ H_2(g) \Longleftrightarrow CH_3OH(l)\ \Delta H° = -128.1$ kJ/mol rxn

The data in the table to the right were determined at 25°C.

(a) Calculate $\Delta G°$ for the reaction above at 25°C.

(b) Calculate K_{eq} for the reaction above at 25°C.

(c) Calculate $\Delta S°$ for the reaction above at 25°C.

(d) In the table above, there is no data for H_2. What are the values of $\Delta H_f°$, $\Delta G_f°$, and of the absolute entropy, S°, for H_2 at 25°C?

Substance	$\Delta H_f°$ (kJ mol^{-1})	$\Delta G_f°$ (kJ mol^{-1})	S° (J mol^{-1} K^{-1})
CO(g)	-110.5	-137.3	+197.9
$CH_3OH(l)$	-238.6	-166.2	+126.8

a. $\Delta G = [-166.2] - [-137.3] = \boxed{-28.9\text{ kJ}}$

b. $\Delta G = -RT \ln K$

$-28900 = (-8.3145)(298) \ln K$

$\ln K = 11.66$

$\boxed{K = 1.16 \times 10^5}$

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

c. $-28.9 = -128.1 - 298\Delta S$

$99.2 = -298\Delta S$

$\boxed{\Delta S = -0.333\text{ kJ/K} \quad \text{or} \quad -333\text{ J/K}}$

d. $\Delta H_f(H_2) = \emptyset$

$\Delta G_f(H_2) = \emptyset$

$\Delta S = \Sigma S_p - \Sigma S_r$

$-333 = [126.8] - [197.9 + 2x]$

$-459.8 = -197.9 - 2x$

$-261.9 = -2x$

$\boxed{x = 131\text{ J/K}}$