Additional Problems: AP Like

$$C(s) + CO_2(g) \rightleftarrows 2\,CO(g)$$

Carbon (graphite), carbon dioxide, and carbon monoxide form an equilibrium mixture, as represented by the equation above.

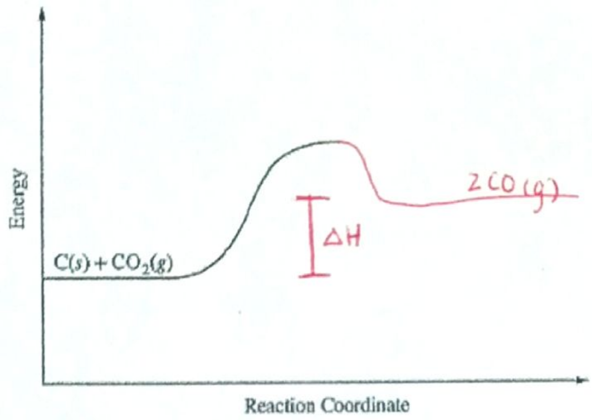
(a) Predict the sign for the change in entropy, ΔS, for the reaction. Justify your prediction. ΔS is positive. 2 moles gas on product side vs. 1 mole gas on reactant side.

(b) In the table below are data that show the percent of CO in the equilibrium mixture at two different temperatures. Predict the sign for the change in enthalpy, ΔH, for the reaction. Justify your prediction.

Temperature	% CO
700°C	60
850°C	94

ΔH is positive. Increasing temperature favors the endothermic process, or forward reaction.

(c) Appropriately complete the potential energy diagram for the reaction by finishing the curve on the graph below. Also, clearly indicate ΔH for the reaction on the graph.



(d) If the initial amount of $C(s)$ were doubled, what would be the effect on the percent of CO in the equilibrium mixture? Justify your answer.

The percent of CO in the equilibrium mixture would remain unchanged because the stoichiometric ratios are fixed.

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 $N_2(g)$ 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.

$2NO(g) + 2CO(g) \rightarrow N_2(g) + 2CO_2(g)$

(b) Write the expression for the equilibrium constant, K_p, for the reaction.

$K_p = \dfrac{P_{N_2}(P_{CO_2})^2}{(P_{NO})^2(P_{CO})^2}$

(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. $\Delta G^\circ = (0 + 2(-394.36)) - 2(86.55) - 2(-137.15) = \boxed{-687.52 \frac{kJ}{mol}}$

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

$\Delta S = -(\Delta G - \Delta H)/T = \boxed{-0.196\ kJ/mol\cdot K}$

(d) For the reaction at 298 K, the value of K_p is 3.3×10^{120}. In an urban area, typical pressures of the gases in the reaction are $P_{NO} = 5.0 \times 10^{-7}$ atm, $P_{CO} = 5.0 \times 10^{-5}$ atm, $P_{N_2} = 0.781$ atm, and $P_{CO_2} = 3.1 \times 10^{-4}$ atm.

$Q = \dfrac{P_{N_2}P_{CO_2}^2}{P_{NO}^2 P_{CO}^2} = \dfrac{(.781)(3.1\times10^{-4})^2}{(5.0\times10^{-7})^2(5.0\times10^{-5})^2} = 1.20 \times 10^{14}$

(i) Calculate the value of ΔG for the reaction at 298 K when the gases are at the partial pressures given above. $\Delta G = \Delta G^\circ + RT\ln Q = -687520 \frac{J}{mol} + (8.314 \frac{J}{mol K})(298K)(\ln 1.20\times10^{14}) = -652.638 \frac{kJ}{mol}$

(ii) In which direction (to the right or to the left) will the reaction be spontaneous at 298 K with these partial pressures? Explain.

$\Delta G < 0$, therefore the reaction will shift right to reach equilibrium.