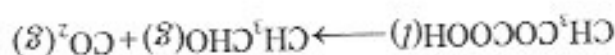


Name: _____

1. The decarboxylation of pyruvic acid occurs via the following reaction:



Given the following thermodynamic data

$$\begin{aligned} \Delta_f H(25^\circ\text{C})_{\text{CH}_3\text{COCOOH}} &= -584 \text{ kJ mol}^{-1} & \Delta_f G(25^\circ\text{C})_{\text{CH}_3\text{COCOOH}} &= -463 \text{ kJ mol}^{-1} \\ \Delta_f H(25^\circ\text{C})_{\text{CH}_3\text{CHO}} &= -166 \text{ kJ mol}^{-1} & \Delta_f G(25^\circ\text{C})_{\text{CH}_3\text{CHO}} &= -133 \text{ kJ mol}^{-1} \\ \Delta_f H(25^\circ\text{C})_{\text{CO}_2} &= -394 \text{ kJ mol}^{-1} & \Delta_f G(25^\circ\text{C})_{\text{CO}_2} &= -394 \text{ kJ mol}^{-1} \end{aligned}$$

a. Calculate $\Delta G_{\text{rxn}}^\circ$. Is this reaction spontaneous under standard state conditions? Justify

your answer.

$$\Delta G_{\text{rxn}}^\circ = \left[(-133 \frac{\text{kJ}}{\text{mol}}) + (-394 \frac{\text{kJ}}{\text{mol}}) \right] - \left[(-463 \frac{\text{kJ}}{\text{mol}}) \right] = -64 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta H_{\text{rxn}}^\circ = \left[(-166 \frac{\text{kJ}}{\text{mol}}) + (-394 \frac{\text{kJ}}{\text{mol}}) \right] - \left[(-584 \frac{\text{kJ}}{\text{mol}}) \right] = 24 \frac{\text{kJ}}{\text{mol}}$$

the reaction is spontaneous under standard conditions because $\Delta G < 0$.

b. Calculate the equilibrium constant, K_p , for this reaction at 80.0 K.

$$\ln \left(\frac{K_2}{K_1} \right) = \frac{-\Delta H_{\text{rxn}}^\circ}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

$$\ln \left(\frac{K_2}{16.045 \times 10^{-12}} \right) = \left(\frac{-0.06344 \text{ kJ/mol}}{8.314 \text{ J/mol K}} \right) \left(\frac{1}{80} - \frac{1}{298} \right)$$

$$K_2 = 2.077 \times 10^{-23}$$

$$K_p = e^{\left(\frac{-64 \text{ kJ/mol}}{8.314 \text{ J/mol K}} \right) \left(\frac{1}{298} \right)} = 6.045 \times 10^{-12}$$

c. At the lower temperature, does the reaction favor the reactants or the products?
at low T the rxn favors products