The conversion of hydroquinone (C₆H₆O₂(aq)) to quinone (C₆H₄O₂(aq)) is involved in many important biochemical reactions. The bombardier beetle, for example, uses the explosive reaction between hydroquinone and hydrogen peroxide (as described by the equation below) as a defence mechanism.

\[ \text{C}_6\text{H}_6\text{O}_2(\text{aq}) + \text{H}_2\text{O}_2(\text{aq}) \rightarrow \text{C}_6\text{H}_4\text{O}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) \]

From the following reaction data, calculate \( \Delta H_{\text{rxn}} \) for the reaction between 1.00 mol of hydroquinone and 1.00 mol of hydrogen peroxide.

\[
\begin{align*}
\text{C}_6\text{H}_6\text{O}_2(\text{aq}) &\rightarrow \text{C}_6\text{H}_4\text{O}_2(\text{aq}) + \text{H}_2(\text{g}) & \Delta H_{\text{rxn}} &= +177.4 \text{ kJ mol}^{-1} \\
\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) &\rightarrow 2\text{H}_2\text{O}_2(\text{aq}) & \Delta H_{\text{rxn}} &= +189.1 \text{ kJ mol}^{-1} \\
\text{H}_2\text{O}(\text{l}) &\rightarrow \text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) & \Delta H_{\text{rxn}} &= +285.8 \text{ kJ mol}^{-1}
\end{align*}
\]

Use the answer you obtained above to calculate the heat liberated (in joules) in the oxidation of 3.86 \( \times \) 10⁻⁴ mol of hydroquinone to quinone.

\[
\Delta H_{\text{rxn}} =
\]

Answer:

Calculate the temperature rise of 0.250 g of water for this quantity of heat.
(The heat capacity of water, \( C_p = 4.184 \text{ J K}^{-1} \text{ g}^{-1} \))

\[
\text{Answer:}
\]
In an experiment, NOCl (2.00 mol) was placed in a closed 1.00 L flask. After equilibrium was established at 25 °C, the concentration of NO(g) was 0.66 M. Calculate the value of $K_c$ at 25 °C for the following reaction.

$$2 \text{NOCl}(g) \rightleftharpoons 2 \text{NO}(g) + \text{Cl}_2(g)$$

$K_c =$

Calculate the value of $K_p$ at 25 °C for the reaction above.

$K_p =$

Given that $\Delta H^\circ_f$ for NOCl(g) = 51.71 kJ mol$^{-1}$ and $\Delta H^\circ_f$ for NO(g) = 90.29 kJ mol$^{-1}$ at 25 °C, calculate the value of $\Delta H^\circ$ for the reaction above.

$\Delta H^\circ_{\text{rxn}} =$

What is the effect upon the [NOCl] of an equilibrium mixture if the temperature is increased?

In which direction will the equilibrium shift if the volume of the flask is reduced?
Equal volumes of carbon monoxide and hydrogen gas are introduced into a sealed 4.5 L flask at 1200 K and the following equilibrium is established.

\[ \text{CO(g)} + 3\text{H}_2(\text{g}) \rightleftharpoons \text{CH}_4(\text{g}) + \text{H}_2\text{O(g)} \quad \Delta H^\circ = -205.9 \text{ kJ mol}^{-1} \]

At equilibrium, the flask contains 0.22 mol of CH\(_4\) and the total pressure in the flask is 46.4 atm. Calculate the amount of H\(_2\)(g) (in mol) that was initially introduced into the flask.

Answer:

In a separate experiment, it is determined that the reaction is in equilibrium when the same 4.5 L flask contains 0.18 mol of CH\(_4\), 0.24 mol of H\(_2\)O, 0.82 mol of CO and 0.65 mol of H\(_2\) at 1200 K. Calculate the concentration equilibrium constant, \(K_c\), for this temperature.

\[ K_c = \]
Calculate the partial pressure equilibrium constant, $K_p$, at 1200 K.

\[ K_p = \]

What is the standard free energy change $\Delta G^\circ$ for the forward reaction (in kJ mol$^{-1}$) at 1200 K?

\[ \Delta G^\circ = \]

What will be the effect on the equilibrium if CO(g) is injected into the flask, which maintains a constant volume.

What will be the effect on the equilibrium if the temperature is decreased?

What will be the effect on the equilibrium if the volume of the flask is decreased?

What will be the effect on the equilibrium if the walls of the flask are refrigerated so that liquid water condenses out?