Calculate the partial pressure equilibrium constant, K_p, at 1200 K.

Marks
7

In the reaction $CO(g) + 3H_2(g) \rightleftharpoons CH_4(g) + H_2O(g)$, 4 moles of gas react to form 2 moles of gas – a *decrease* of 2 moles of gas or Δn = -2 mol.

From the previous question (2004-N-4), K_c = 3.9. As K_p and K_c are related by $K_p = K_c(RT)^{\Delta n}$

$$K_p = 3.9 \times (0.08206 \times 1200)^{-2} = 4.0 \times 10^{-4}$$

K_p = **4.0 × 10^{-4}**

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

Using $\Delta G° = -RT\ln K_p$:

$$\Delta G° = -(8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (1200 \text{ K}) \times \ln(4.0 \times 10^{-4})$$
$$= -78000 \text{ J mol}^{-1} = -78 \text{ kJ mol}^{-1}$$

$\Delta G°$= **-78 kJ mol^{-1}**

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

The equilibrium will shift to reduce the amount of CO(g). It will shift towards products.

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

The reaction is exothermic (see previous question (2004-N-4)) so the reaction will shift to give out more heat. It will shift towards products.

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

If the volume is decreased, the pressure will increase. The reaction will shift to reduce the pressure. As the forward reaction leads to a decrease in the amount of gas, the equilibrium will shift towards products.

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

If $H_2O(g)$ is removed to form pure $H_2O(l)$, the equilibrium will shift to increase the concentration of $H_2O(g)$. The equilibrium will shift towards products.