Marks 4

The CO(g) in water gas can be reacted further with $H_2O(g)$ in the so-called "water-gas shift" reaction:

$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$$

At 900 K, $K_c = 1.56$ for this reaction. A sample of water gas flowing over coal at 900 K contains a 1:1 mole ratio of CO(g) and $H_2(g)$, as well as 0.250 mol L^{-1} $H_2O(g)$. This sample is placed in a sealed container at 900 K and allowed to come to equilibrium, at which point it contains 0.070 mol L^{-1} $CO_2(g)$. What was the initial concentration of CO(g) and $H_2(g)$ in the sample?

The reaction table is

	CO(g)	**$H_2O(g)$**	$\rightleftharpoons$	**$CO_2(g)$**	**$H_2(g)$**
initial	**x**	**0.250**		**0**	**x**
change	**-0.070**	**-0.070**		**+0.070**	**+0.070**
equilibrium	**x – 0.070**	**0.250 – 0.070**		**0.070**	**x + 0.070**

The equilibrium constant in terms of concentrations, K_c, is:

$$\mathbf{K_c = \frac{[CO_2(g)][H_2(g)]}{[H_2O(g)][CO(g)]} = \frac{(0.070)(x + 0.070)}{(0.180)(x - 0.070)} = 1.56}$$

$$\mathbf{x = [CO(g)]_{initial} = [H_2(g)]_{initial} = 0.12\ mol\ L^{-1}}$$

$[CO] = [H_2] =$ **$0.12\ mol\ L^{-1}$**

If the walls of the container are chilled to below 100 ºC, what will be the effect on the concentration of $CO_2(g)$?

At temperatures below 100 °C, the water vapour will condense to form $H_2O(l)$. Following Le Chatelier's principle, the equilibrium will shift to the left as $[H_2O(g)]$ is reduced by this process and so $[CO_2(g)]$ will decrease.