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:

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

$x = [CO(g)]_{initial} = [H_2(g)]_{initial} = 0.12$ 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.