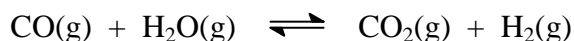


The CO(g) in water gas can be reacted further with H₂O(g) in the so-called “water-gas shift” reaction:



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 H₂(g), as well as 0.250 mol L⁻¹ H₂O(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 CO₂(g). What was the initial concentration of CO(g) and H₂(g) in the sample?

Marks
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The reaction table is

	CO(g)	H ₂ O(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:

$$K_c = \frac{[\text{CO}_2\text{(g)}][\text{H}_2\text{(g)}]}{[\text{H}_2\text{O(g)}][\text{CO(g)}]} = \frac{(0.070)(x + 0.070)}{(0.180)(x - 0.070)} = 1.56$$

$$x = [\text{CO(g)}]_{\text{initial}} = [\text{H}_2\text{(g)}]_{\text{initial}} = 0.12 \text{ mol L}^{-1}$$

$$[\text{CO}] = [\text{H}_2] = 0.12 \text{ mol L}^{-1}$$

If the walls of the container are chilled to below 100 °C, what will be the effect on the concentration of CO₂(g)?

At temperatures below 100 °C, the water vapour will condense to form H₂O(l). Following Le Chatelier’s principle, the equilibrium will shift to the left as [H₂O(g)] is reduced by this process and so [CO₂(g)] will decrease.