Marks 5

- Methane, CH_4, reacts with hydrogen sulfide, H_2S, according the following equilibrium:

$$CH_4(g) + 2H_2S(g) \rightleftharpoons CS_2(g) + 4H_2(g)$$

In an experiment 1.00 mol of CH_4, 2.00 mol of H_2S, 1.00 mol of CS_2 and 2.00 mol of H_2 are mixed in a 250 mL vessel at 960 °C. At this temperature, $K_c = 0.034$ (based on a standard state of 1 mol L^{-1}).

Calculate the reaction quotient, Q, and hence predict in which direction the reaction will proceed to reach equilibrium? Explain your answer.

Using concentration = number of moles / volume, the concentrations when the gases are mixed are:

$[CH_4(g)] = 1.00 \text{ mol} / 0.250 \text{ L} = 4.00 \text{ mol L}^{-1}$
$[H_2S(g)] = 2.00 \text{ mol} / 0.250 \text{ L} = 8.00 \text{ mol L}^{-1}$
$[CS_2(g)] = 1.00 \text{ mol} / 0.250 \text{ L} = 4.00 \text{ mol L}^{-1}$
$[H_2(g)] = 2.00 \text{ mol} / 0.250 \text{ L} = 8.00 \text{ mol L}^{-1}$

From the chemical equation, the reaction quotient is:

$$Q = \frac{[CS_2(g)][H_2(g)]^4}{[CH_4(g)][H_2S(g)]^2} = \frac{(4.00)(8.00)^4}{(4.00)(8.00)^2} = 64.0$$

As $Q > K_c$, therefore the reaction will shift to the left until $Q = K_c$

Show that the system is at equilibrium when $[CH_4(g)] = 5.56$ M.

A reaction table can be constructed to calculate the equilibrium concentrations:

	$CH_4(g)$ +	$2H_2S(g)$	$\rightleftharpoons$	$CS_2(g)$ +	$4H_2(g)$
Initial	**4.00**	**8.00**		**4.00**	**8.00**
Change	**$+x$**	**$+2x$**		**$-x$**	**$-4x$**
Equilibrium	**$4.00 + x$**	**$8.00 + 2x$**		**$4.00 - x$**	**$8.00 - 4x$**

If $[CH_4(g)]_{equilibrium} = 5.56$ M then $4.00 + x = 5.56$ M and $x = 1.56$ M. Hence:

$[CH_4(g)]_{equilibrium} = (4.00 + x)$ M = 5.56 M
$[H_2S(g)]_{equilibrium} = (8.00 + 2x)$ M = 11.12 M
$[CS_2(g)]_{equilibrium} = (4.00 - x)$ M = 2.44 M
$[H_2(g)]_{equilibrium} = (8.00 - 4x)$ M = 1.76 M.

With these concentrations:

$$K_c = \frac{[CS_2(g)][H_2(g)]^4}{[CH_4(g)][H_2S(g)]^2} = \frac{(2.44)(1.76)^4}{(5.56)(11.12)^2} = 0.034$$