Q5 Chemical Equilibrium **[3 + 4 = 7 marks]**

a) Consider the equilibrium:

$$CH_4(g) + 2\,H_2S(g) \rightleftharpoons CS_2(g) + 4\,H_2(g) \qquad \Delta_r H^\circ \text{ is positive}$$

In which direction will the equilibrium be shifted by the following changes? Explain.

1 mark each (½ for answer, ½ for explanation) [3 marks]

(i) addition of $CH_4(g)$

Increasing [CH_4] instantaneously decreases the value of *Q* by increasing the value of the bottom line of the equilibrium quotient expression. Therefore, the system shifts to the right to consume some of the added CH_4, thereby increasing the value of *Q*.

(ii) increasing the temperature of the reaction mixture

As the forward reaction is endothermic, *K* increases as the temperature increases. *Q* therefore increases in order to maintain equilibrium and so the system shifts to the right.

(iii) decreasing the volume of the container at constant temperature

***Q* is proportional to $1/V^2$ and so decreasing V instantaneously increases the value of *Q*. The system therefore shifts to the left, thereby decreasing the value of *Q*. The side with fewer gas molecules is favoured.**

Q5 (continued) Chemical Equilibrium **[3 + 4 = 7 marks]**

b) $K_c = 3.2 \times 10^{-34}$ for the following reaction at 25 °C

$$2\ HCl(g) \rightleftharpoons H_2(g) + Cl_2(g)$$

The reaction vessel initially contained 0.0500 mol L^{-1} HCl.
Calculate the concentrations of H_2 and Cl_2 at equilibrium.

[4 marks]

The balanced chemical equation is: 2 HCl(g) $\rightleftharpoons$ $H_2(g) + Cl_2(g)$, so

$$K_c = \frac{[H_2][Cl_2]}{[HCl]^2} = 3.2 \times 10^{-34}$$

The concentration table is: **1 mark**

	[HCl]	[H_2]	[Cl_2]
Initial	0.0500	0	0
Change	$-2x$	$+x$	$+x$
Equilibrium	$0.0500-2x$	$+x$	$+x$

Substituting these equilibrium concentrations into the equilibrium constant expression gives:

$$K_c = \frac{[H_2][Cl_2]}{[HCl]^2}$$

$$= \frac{(x)(x)}{(0.0500 - 2x)^2}$$ **1 mark**

$$= 3.2 \times 10^{-34}$$

Since K_c is small, we can assume that $(0.0500 - 2x) \approx 0.0500$. Thus:

$$\frac{x^2}{(0.0500)^2} = 3.2 \times 10^{-34}$$ **1 mark**

So: $x = 8.9 \times 10^{-19}$ M = [H_2] = [Cl_2]
[HCl] = $(0.0500 - x)$ ≈ 0.0500 M

Concentrations of H_2 and Cl_2 at equilibrium will each be 8.9×10^{-19} mol L^{-1} **1 mark**

(or calculation making the assumptions without using the table)
Answer should be 2 significant figures.