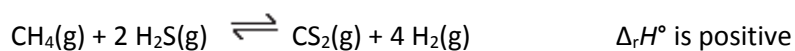


Q5 Chemical Equilibrium

[3 + 4 = 7 marks]

a) Consider the equilibrium:



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 $\text{CH}_4(\text{g})$

Increasing $[\text{CH}_4]$ 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.

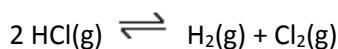
(iii) decreasing the volume of the container at constant temperature

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



The reaction vessel initially contained 0.0500 mol L⁻¹ HCl.
Calculate the concentrations of H₂ and Cl₂ at equilibrium.

[4 marks]

The balanced chemical equation is: $2 \text{HCl(g)} \rightleftharpoons \text{H}_2\text{(g)} + \text{Cl}_2\text{(g)}$, so

$$K_c = \frac{[\text{H}_2][\text{Cl}_2]}{[\text{HCl}]^2} = 3.2 \times 10^{-34}$$

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:

$$\begin{aligned} K_c &= \frac{[\text{H}_2][\text{Cl}_2]}{[\text{HCl}]^2} \\ &= \frac{(x)(x)}{(0.0500 - 2x)^2} \\ &= 3.2 \times 10^{-34} \end{aligned}$$

1 mark

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

$$\begin{aligned} \text{So: } x &= 8.9 \times 10^{-19} \text{ M} = [\text{H}_2] = [\text{Cl}_2] \\ [\text{HCl}] &= (0.0500 - x) \approx 0.0500 \text{ M} \end{aligned}$$

Concentrations of H₂ and Cl₂ 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.