

Answers to exam-style questions

Topic 7

- 1 C
- 2 B
- 3 A
- 4 A
- 5 C
- 6 A
- 7 D
- 8 B
- 9 C
- 10 D

11 a In dynamic equilibrium, macroscopic properties are constant/the concentrations of all reactants and products remain constant. The rate of the forward reaction is equal to the rate of the reverse reaction. [2]

b i The position of equilibrium shifts to the right because the forward reaction is endothermic. As the temperature is increased, the position of equilibrium shifts in the direction that minimises the effect of the change – to the right (endothermic direction) – to take in heat. The value of the equilibrium constant increases because more H_2O and CO are present at equilibrium. The expression for the equilibrium constant is:

$$K_c = \frac{[\text{H}_2\text{O}][\text{CO}]}{[\text{H}_2][\text{CO}_2]}$$

The concentrations of H_2O and CO are on the top of this expression. [3]

ii The position of equilibrium does not change; there is the same number of moles of gas on both sides of the equation. The value of the equilibrium constant does not change; only a change in temperature affects the value of the equilibrium constant. [2]

iii The position of equilibrium and the value of the equilibrium constant do not change; a catalyst speeds up the forward and reverse reactions equally. [2]

12 a $K_c = \frac{[\text{NO}(\text{g})]^2}{[\text{N}_2(\text{g})][\text{O}_2(\text{g})]}$ [1]

b As the temperature increases, the value of the equilibrium constant increases. This means that the position of equilibrium has shifted to the right (more NO present at equilibrium). The forward reaction is therefore endothermic; when the temperature increases the position of equilibrium shifts in the endothermic direction to take in heat and minimise the effect of the change. [3]

13 a $K_c = \frac{[\text{NO}_2(\text{g})]^2}{[\text{N}_2\text{O}_4(\text{g})]}$ [1]

b The mixture will become browner. The forward reaction is endothermic, and as the temperature is increased, the position of equilibrium shifts in the endothermic direction (to the right) to take in heat and minimise the effect of the change, so there will be more NO_2 (brown) in the equilibrium mixture. [3]

c The equilibrium mixture will be less brown/more colourless at higher pressure. There are fewer moles of gas on the left-hand side and as the pressure is increased, the position of equilibrium will shift to the left to minimise the effect of the change. More N_2O_4 (colourless) will be present at equilibrium. The value of the equilibrium constant will not change. Only a change in temperature affects the value of the equilibrium constant. [4]



initial no. moles/mol 0.200 0.00

equilibrium no. moles/mol 0.140

The number of moles of N_2O_4 that has reacted is $0.200 - 0.140 = 0.060$ mol. From the chemical equation we can see that 1 mol N_2O_4 reacts to form 2 mol NO_2 . Therefore 0.060 mol N_2O_4 react to form $2 \times 0.060 = 0.120$ mol NO_2 , so 0.120 mol NO_2 are present at equilibrium.

The volume of the container is 4.00 dm^3 .

Equilibrium concentrations are worked out using:

concentration = no. moles/volume

Therefore the equilibrium concentrations are:

$$[\text{N}_2\text{O}_4] = \frac{0.140}{4.00} = 0.0350 \text{ mol dm}^{-3}$$

$$[\text{NO}_2] = \frac{0.120}{4.00} = 0.0300 \text{ mol dm}^{-3}$$

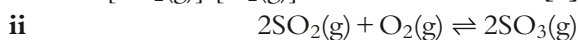
The expression for K_c is:

$$K_c = \frac{[\text{NO}_2(\text{g})]^2}{[\text{N}_2\text{O}_4(\text{g})]}$$

The values can be put into this expression:

$$K_c = \frac{0.0300^2}{0.0350} = 0.0257 \quad [4]$$

14 a i $K_c = \frac{[\text{SO}_3(\text{g})]^2}{[\text{SO}_2(\text{g})]^2[\text{O}_2(\text{g})]} \quad [1]$



initial no. moles/ 0.120 0.120 0.00

mol

equilibrium 0.060

no. moles/mol

From the chemical equation we can work out that 0.030 mol O_2 must have reacted to produce 0.060 mol SO_3 (the mole ratio is 1:2 in the equation). Therefore the number of moles of O_2 left at equilibrium is $0.120 - 0.030 = 0.090$ mol.

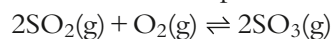
The volume of the reaction vessel is 2.00 dm^3 .

$$\text{concentration} = \frac{\text{no. moles}}{\text{volume}}$$

Therefore:

$$\begin{aligned} \text{concentration of } \text{O}_2 \text{ at equilibrium} &= \frac{0.090}{2.00} \\ &= 0.045 \text{ mol dm}^{-3} \end{aligned} \quad [2]$$

iii To work out the equilibrium constant, we need the equilibrium concentrations of all species.



initial no. moles/ 0.120 0.120 0.00

mol

equilibrium no. 0.090 0.060

moles/mol

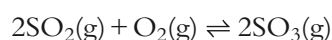
From the chemical equation we can work out that 0.060 mol SO_2 must have reacted to produce 0.060 mol SO_3 (the mole ratio is 2:2 in the equation). Therefore the number of moles of SO_2 left at equilibrium is $0.120 - 0.060 = 0.060$ mol.

The volume of the reaction vessel is 2.00 dm^3 .

$$\text{concentration} = \frac{\text{no. moles}}{\text{volume}}$$

Therefore:

$$\begin{aligned} \text{concentration of } \text{SO}_2 \text{ at equilibrium} &= \frac{0.060}{2.00} \\ &= 0.030 \text{ mol dm}^{-3} \end{aligned}$$



initial no. moles/ 0.120 0.120 0.00

mol

equilibrium no. 0.060 0.090 0.060

moles/mol

equilibrium 0.030 0.045 0.030

concentration/ mol dm^{-3}

$$K_c = \frac{[\text{SO}_3(\text{g})]^2}{[\text{SO}_2(\text{g})]^2[\text{O}_2(\text{g})]}$$

Substituting in the equilibrium concentrations we get:

$$K_c = \frac{0.030^2}{0.030^2 \times 0.045} = 22 \quad [3]$$



We have been given enthalpy change of formation values and can therefore use the equation:

$$\Delta H_r = \Sigma \Delta H_f(\text{products}) - \Sigma \Delta H_f(\text{reactants})$$

To calculate the enthalpy change for the reaction.

$$\Sigma \Delta H_f(\text{products}) = 2 \times -395 = -790 \text{ kJ mol}^{-1}$$

$$\Sigma \Delta H_f(\text{reactants}) = 2 \times -297 = -594 \text{ kJ mol}^{-1}$$

The enthalpy change of formation of $\text{O}_2(\text{g})$ is zero, as it is an element in its standard state.

$$\Delta H_r = -790 - (-594) = -196 \text{ kJ mol}^{-1} \quad [2]$$

ii The reaction is exothermic in the forward direction; so as the temperature is increased from 1100 K to 1500 K the position of equilibrium shifts in the endothermic direction, i.e. to the left. There will be less SO_3 present at equilibrium and because SO_3 appears on the top of the K_c expression, the value of K_c will be lower at 1500 K. **[2]**

15 a $K_c = \frac{[\text{HI}(\text{g})]^2}{[\text{H}_2(\text{g})][\text{I}_2(\text{g})]} \quad [1]$

b The expression for the equilibrium constant for this reaction would be:

$$K_c = \frac{[\text{H}_2(\text{g})][\text{I}_2(\text{g})]}{[\text{HI}(\text{g})]^2}$$

This is the inverse of the equilibrium constant expression in **a** so the value of the

$$\text{equilibrium constant is } \frac{1}{54.0}, \text{ i.e. } 0.0185. \quad [1]$$



initial no. moles/mol 0.00 0.00 2.00

equilibrium x x $2 - 2x$

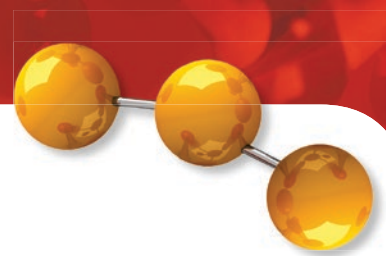
no. moles/mol

equilibrium x x $(2 - 2x)$

concentration/ mol dm^{-3} $\frac{x}{4.00}$ $\frac{x}{4.00}$ $\frac{(2 - 2x)}{4.00}$

We will assume that when the HI comes to equilibrium $2x$ moles react. From the chemical equation we can see that $2x$ moles of HI will form x moles of H_2 and x moles of I_2 .

$$K_c = \frac{[\text{HI}(\text{g})]^2}{[\text{H}_2(\text{g})][\text{I}_2(\text{g})]}$$



Substituting the equilibrium concentrations
we get:

$$K_c = \frac{\left[\frac{(2-2x)}{4.00}\right]^2}{\left(\frac{x}{4.00}\right)\left(\frac{x}{4.00}\right)}$$

Because the number of moles is the same on both sides of the chemical equation (two moles of HI and a total of two moles of H₂ and I₂), the volumes cancel out:

$$K_c = \frac{(2-2x)^2}{(x)(x)}$$

We are given that $K_c = 54.0$, so we can write:

$$54.0 = \frac{(2-2x)^2}{x^2}$$

The right-hand side is a perfect square, and taking the square root of both sides we get:

$$\sqrt{54.0} = \frac{2-2x}{x}$$

$$7.39 = \frac{2-2x}{x}$$

Rearranging this we get:

$$7.39x = 2 - 2x$$

$$9.39x = 2$$

$$x = \frac{2}{9.39} = 0.214 \text{ mol}$$

This is the equilibrium number of moles of H₂, and to determine the equilibrium concentration we must divide by the volume:

$$\begin{aligned} \text{equilibrium concentration of H}_2 &= \frac{0.214}{4.00} \\ &= 0.0535 \text{ mol dm}^{-3} \end{aligned}$$

[3]