09 Chemical Equilibria

Set B **Answer key**

A Multiple choice

4 MARKS

1

B The addition of a catalyst lowers the activation energy of the forward reaction but not that of the reverse reaction. [1]

A catalyst speeds up the rate of a reaction by providing an alternative pathway with a lower activation energy (A is correct). A catalysed and an uncatalysed reaction have the same enthalpy change (C is correct).

The activation energies for both the forward and backward reactions are lowered by the same amount (B is incorrect). Hence, a catalyst does not change the position of equilibrium (D is correct). Instead, it shortens the time to reach equilibrium.

2

B 8.27×10^{-5} [1]

At equilibrium, $[NH_3] = [HNO_3] = \frac{0.0113}{0.83} \times \frac{66.8}{100} = 0.00909$ mol dm^{-3}

$K_c = [NH_3][HNO_3] = (0.00909)(0.00909) = 8.27 \times 10^{-5}$ mol^2 dm^{-6}

3

C 2 and 3 only [1]

1 Incorrect. This is an exothermic reaction, hence increasing the temperature will shift the position of equilibrium to the left, favouring the endothermic backward reaction to absorb the additional heat. Fewer products, smaller equilibrium constant.
2 Correct. The rate constant for the forward reaction increases as temperature increases.
3 Correct. The rate constant for the backward reaction increases as temperature increases.
4 Incorrect. The activation energy for the forward reaction does not change with temperature (activation energy only decreases when a catalyst is introduced).

4

B $2P \rightleftharpoons 2Q + R$ [1]

Let a and b be the stoichiometric coefficients of P and Q: $aP \rightleftharpoons bQ + R$. Equilibrium amounts: P (2 – ax), Q bx, R x.

At equilibrium, total amount = 2 – ax + bx + x = 2 + x, so 2 – ax + bx = 2. Therefore a = b, i.e. P and Q have the same stoichiometric coefficients.

B Structured questions

23 MARKS

5 (a)

$K_c = \dfrac{[CO_2][H_2]}{[CO][H_2O]}$ [1]

NOTE $[H_2O]$ must be included, as this is a homogeneous system: all the species are gases.

(b)

	CO(g)	$H_2O(g)$	$CO_2(g)$	$H_2(g)$
initial / mol	0.200	0.200	0	0
change / mol	−0.2 × 0.775	−0.2 × 0.775	+0.2 × 0.775	+0.2 × 0.775
equilibrium / mol	0.045	0.045	0.155	0.155

$K_c = \dfrac{(0.155)^2}{(0.045)^2}$ [1]

$= 11.86 \approx \mathbf{11.9}$ [1]

(c)

$-18.4 \times 10^3 = -(8.31)(600) \ln K_c$

$\ln K_c = 3.690$, so $K_c = 40.04 \approx \mathbf{40.0}$ [1]

NOTE Convert ΔG to J mol^{-1} first.

6 (a)

$\Delta H = 9.2 - 2(+33.2) = \mathbf{-57.2\ kJ\ mol^{-1}}$ [1]

(b)

The dimerisation will be favoured by high pressure. At high pressure, the position of equilibrium will shift to decrease pressure by decreasing the total number of gaseous molecules present. Since there are two molecules of gas on the left-hand side of the equation and one on the right, the forward reaction is favoured, and the position of equilibrium shifts to the right to form more N_2O_4. [1]

The dimerisation will be favoured by low temperature. At low temperature, the position of equilibrium will shift to increase temperature. Since the forward reaction is exothermic, the forward reaction is favoured, and the position of equilibrium shifts to the right to form more N_2O_4. [1]

(c)

Tube in iced water: pale yellow. Tube in hot water: brown. [1]

In iced water, the low temperature favours the exothermic forward reaction, so the position of equilibrium shifts to the right to form more colourless N_2O_4. In hot water, the high temperature favours the endothermic backward reaction, so the position of equilibrium shifts to the left to form more brown NO_2. [1]

7 (a)

Equilibrium amounts / mol: H_2O 4 − x, CO 4 − x, CO_2 x, H_2 x. Let the volume of the vessel be V dm^3.

$K_c = \dfrac{(x/V)^2}{((4-x)/V)^2} = 9$ [1]

Solving, x = 3 mol. Hence, the equilibrium amount of CO = **1 mol** [1]

(b) $\chi_{H_2} = \chi_{CO_2} = \frac{3}{8}$, $\chi_{H_2O} = \chi_{CO} = \frac{1}{8}$ [1]

(c) (i) At a lower temperature, the rate at which equilibrium is attained is slower. By Le Chatelier's Principle, since $\Delta H < 0$, the position of equilibrium shifts to the right and produces heat so as to increase the temperature. Hence, the amount of CO_2 is higher at equilibrium. K_c is also increased. [1]

(ii) A decrease in volume, which leads to an increase in concentration, favours neither the backward nor the forward reaction, since the number of molecules on both sides of the reaction is the same. The amount of CO_2 produced is the same. The rate at which equilibrium is attained increases. [1]

(iii) In the presence of a suitable catalyst, the rate at which equilibrium is attained is faster. However, there is no effect on the amount of CO_2 formed, and K_c remains the same. [1]

8 Advantage: by Le Chatelier's principle, a lower temperature favours the forward exothermic reaction, and hence the yield of ammonia increases. [1]

Disadvantage: if the temperature is too low, the reaction is too slow. [1]

9 **(a)(i)** $\Delta H_r = -396 + 297 =$ **-99.0 kJ mol^{-1}** [1]

(ii) $\Delta G = \Delta H - T\Delta S = -99 \times 1000 - (298)(-93.5) =$ **-71.1 kJ mol^{-1}** [1]

(b)(i) $\Delta G = -2.303RT \log K_c$, so $\log K_c = (-71.14 \times 1000) \div (-2.303 \times 8.31 \times 298) = 12.47$ [1]

$K_c = 10^{12.47} =$ **2.95×10^{12} mol$^{-½}$ dm$^{3/2}$** [1]

NOTE R is in J K^{-1} mol^{-1}, so ΔG must be in J mol^{-1}, not kJ mol^{-1}.

(ii) Since $K_c \gg 1$, the position of equilibrium lies to the far right. Hence, the reaction goes almost to completion. [1]

(iii) Although SO_3 is energetically more stable than SO_2, the oxidation of SO_2 to SO_3 is a slow process with a high activation energy. [1]