# Suggested Answers VCE Chemistry Equilibrium Test Unit 3

SECTION A [1 mark per question.]

Q1 A The activation energy for the forward reaction is the energy required to break the bonds necessary for the reaction to occur. In this case it is the bonds between the atoms in the N_2 and O_2 molecules.

Q2 C For an endothermic reaction the energy of the products is greater than the energy of the reactants, as can be seen in the energy profile diagram. Therefore the energy required to break the bonds in the reactants (I) is greater than the energy release in forming the bonds in the products (II).
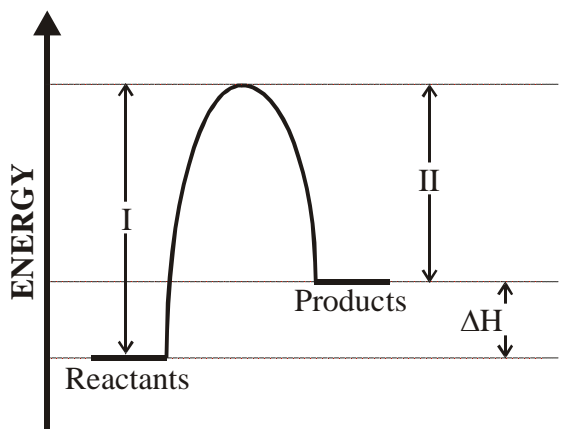
Q3 D The final concentration of hydrogen chloride in the system is 0.12 M.

	$H_2(g)$ +	$Cl_2(g)$	$\rightleftharpoons$	$2HCl(g)$
Stoichiometry		1 mole 1 mole		2 mole
Initial concentrations		0.15 M 0.15 M		0 M
Change in concentrations	0.06 M 0.06 M $\Leftarrow$			0.12 M
Equilibrium concentrations		0.09 M 0.09 M		0.12 M

The value for the equilibrium constant, K, is given by

$$K = \frac{[HCl]^2}{[H_2][Cl_2]} = \frac{(0.12)^2}{(0.09)(0.09)} = \mathbf{1.8}$$

Q4 C The expected pH of pure water is 7.0 at 25 °C, thus the observed pH is higher then expected. Both responses B and D could result in higher pH readings.
Response A states that the ionisation of water is endothermic and that the measurement was made below 25 °C. Lowering the temperature for an endothermic reaction will result in an equilibrium position shift to the left (LeChatelier's principle) lowering the $[H^+]$ and increasing the pH.
Response C states that the ionisation of water is exothermic and that the measurement was made below 25 °C. **Lowering the temperature for an exothermic reaction will result in an equilibrium position shift to the right (LeChatelier's principle) increasing the $[H^+]$ and lowering the pH.**

Q5 D The graph shows that adding a trace amount of aqueous copper(II) sulfate solution increased the rate of reaction. Therefore this solution or one or other of the ions is most likely acting as a catalyst. A catalyst by definition lowers the activation energy for both the forward and reverse reactions by providing an alternative pathway at lower activation energy.

Q6 D The aim is to increase the **equilibrium yield**, to do this the position of equilibrium must be shifted to the right. This can be achieved using LeChatelier's principle by decreasing the temperature as the forward reaction is exothermic, increasing the amount (concentration) of one of the reactants, or increasing the pressure in this case since the products have the lesser number of particles. **Adding a catalyst does not increase the equilibrium yield** only the rate at which the reaction reaches equilibrium.

Q7 A $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$
$n(OH^-) = c \times V = 0.200 \times (50.00/1000) = 0.0100$ mol
This number of mole of $OH^-(aq)$ ions is then diluted with water to 10.0 L
$c(OH^-) = n / V = 0.0100/10.0 = 1.00 \times 10^{-3}$ M
Assuming 25 °C: $K_w = [H^+][OH^-] = 10^{-14} M^2$
$[H^+] = 10^{-14} / [OH^-] = 10^{-14}/1.00 \times 10^{-3} = 10^{-11}$ M
$pH = -\log_{10}[H^+] = -\log_{10}(10^{-11}) = \mathbf{11}$

Q8 B $H_2O(l) + CO_2(g) \rightleftharpoons H_2CO_3(aq) \rightleftharpoons HCO_3^-(aq) + H^+(aq)$

In the lungs carbon dioxide is removed from the blood serum, therefore the concentration of $CO_2(g)$ decreases therefore the concentration of $H_2CO_3(aq)$ will also decrease. Using LeChatelier's principle this will result in a shift in the position of equilibrium to the left therefore decreasing the $[H^+]$ and increasing the pH.

Q9 D The flour dust has a very large surface area therefore this will increase the rate of reaction between the flour and the oxygen. Dust explosions have caused considerable damage to flour mills, wheat silos and coal bunkers over time.

Q10 A Benzoic acid is a weak acid therefore will not fully ionise in water.

$C_6H_5COOH(aq) + H_2O(l) \rightleftharpoons C_6H_5COO^-(aq) + H_3O^+(aq)$

Complete ionisation would result in $[H^+] = 0.010$ M $\Rightarrow$ pH = 2
In this case $[H^+] < 0.010$ M $\Rightarrow$ pH > 2
A pH = 6.8 is too close to neutral.

Q11 C The activation energy for the reverse reaction is given by II.
The energy change for the reaction is the difference between the energy of the reactants and the products and is given by V.

Q12 D When a reaction commences the rate of the forward reaction is much higher that that of the reverse reaction because the concentrations of the reactants are much higher. As the reaction proceeds the rate of the forward reaction decreases and the rate of the reverse reaction increases. This occurs until equilibrium is reached where the rate of the forward and reverse reactions are equal, so as to maintain the constant concentrations of reactants and products.

SECTION B

Question 1 - [9 marks, 11 minutes]
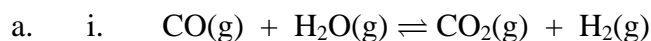
a. i. $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$

$K = \dfrac{[CO_2][H_2]}{[CO][H_2O]}$ **[1 mark]**

  ii. The volume of the flask was 20.0 L $\Rightarrow$ c = n / V

	$CO(g)$ +	$H_2O(g)$	$\rightleftharpoons CO_2(g)$ +	$H_2(g)$
Initially:	1.0 mol	1.0 mol	0 mol	0 mol
	0.050 M	0.050 M	0 M	0 M
Equilibrium information:				0.036 M
Change:	-0.036 M	-0.036 M	$\Leftarrow$ 0.036 M	0.036 M
At equilibrium:	0.014 M	0.014 M	0.036 M	0.036 M

$K = \dfrac{[CO_2][H_2]}{[CO][H_2O]} = \dfrac{(0.036)(0.036)}{(0.014)(0.014)} =$ **6.6 [3 marks]**

b. i. Because **the value of the equilibrium constant is very small, then the extent of reaction is minimal**. Therefore the concentration of chlorine atoms in a sample of chlorine at 298 K would be very low compared with the concentration of chlorine molecules. **[1 mark]**

  ii. c = n / V
Using the general gas equation PV = nRT $\Rightarrow$ c = n/V = P / RT
c = 101.3 /(8.31 × 298) = 0.041 M
Using the molar volume $V_M = 24.5$ L mol^{-1}
c = 1 / V_M = 1 / 24.5 = **0.041 M [1 mark]**

  iii. Let the concentration of atomic chlorine be x M, and since the extent of reaction is minimal then the concentration of chlorine molecules will be unchanged.
[Cl] = x M   [Cl_2] = 0.041 M

$K = \dfrac{[Cl]^2}{[Cl_2]} = 1.4 \times 10^{-38}$ M **[1 mark]**

$\dfrac{x^2}{0.041} = 1.4 \times 10^{-38} \Rightarrow x^2 = 1.4 \times 10^{-38} \times 0.041 = 5.7 \times 10^{-40}$ M

$x = \sqrt{5.7 \times 10^{-40}} = 2.4 \times 10^{-20}$ M

[Cl] = **2.4×10^{-20} M [1 mark]**

  iv. For the reverse reaction the equilibrium constant would be given by the expression,

$K = \dfrac{[Cl_2]}{[Cl]^2}$, which is the reciprocal of the previous equilibrium constant,

therefore
K = 1/1.4×10^{-38} = **7.1×10^{37} M^{-1} [1 mark]**

Question 2 - [5 marks, 6 minutes]

a.   **$ClC_6H_5OH(aq) + H_2O(l) \rightleftharpoons ClC_6H_5O^-(aq) + H_3O^+(aq)$ [1 mark]**

b. $K_a = \dfrac{[ClC_6H_5O^-][H_3O^+]}{[ClC_6H_5OH]}$ **[1 mark]**

The dissociation constant for weak acids do not include the concentration of water, because the change in this due to the reaction of the acid with water is exceptionally small and can essentially be treated as a constant.

c.   pH = 5.3, therefore
$[H_3O^+] = 10^{-pH} = 10^{-5.3} = 5.0 \times 10^{-6}$ M **[1 mark]**
For each $H_3O^+(aq)$ ion formed $ClC_6H_5O^-(aq)$ ions also forms, therefore
$[ClC_6H_5O^-] = [H_3O^+] = 5.0 \times 10^{-6}$ M
Let $[ClC_6H_5OH] = x$, thus

$$K_a = 2.95 \times 10^{-9} = \frac{(5.0 \times 10^{-6})(5.0 \times 10^{-6})}{x}$$

$$x = \frac{(5.0 \times 10^{-6})(5.0 \times 10^{-6})}{2.95 \times 10^{-9}} = \textbf{0.0085 M } \textbf{[1 mark]}$$

d.   Since the pH is 5.3 then $[H_3O^+] = 10^{-pH} = 10^{-5.3} = 5.0 \times 10^{-6}$ M
$K_w = [H_3O^+][OH^-] = 10^{-14}$ M^2 at 25 °C, then
$[OH^-] = 10^{-14} / [H_3O^+] = 10^{-14} / 5.0 \times 10^{-6} = \textbf{2.0} \times \textbf{10}^{\textbf{-9}}$ **M [1 mark]**

Question 3 - [9 marks, 11 minutes]

a.   The system reached equilibrium **4 times**, as can be seen from the different selections of the graph when the concentrations of the two gases remains constant. **[1 mark]**

b.   i.   At t_1 the concentrations of both gases doubles this is due to either **halving the volume, or doubling the pressure, of the system. [1 mark]**

       ii.   Using LeChatelier's principle that a system will shift its position to resist the change made to it. In order to reduce the concentration increase the system will **shift its position of equilibrium to favour the side with the lesser number of particles**. In this case this will be a shift to the left, increasing the concentration of the N_2O_4 molecules. **[1 mark]**

c.   i.   At t_2 the only change is to the NO_2 concentration, therefore **some NO_2 molecules have been removed. [1 mark]**

       ii.   Again using LeChatelier's principle the number of NO_2 particles has been reduced so in order to restore this balance the **position of equilibrium will shift to the right as this will produce $2NO_2$ molecules for each N_2O_4 molecule that breaks down. [1 mark]**

d.   i.   At t_3 the concentration of NO_2 decreases while that of N_2O_4 increases, this is the result of a temperature change to the system. As the **forward reaction is endothermic** and the concentration of the N_2O_4 is increasing this indicates that the **temperature has decreased. [1 mark]** (This is supported by the longer time for the system to re-establish equilibrium.)

       ii.   Using LeChatelier's principle again as the temperature is lowered then the system will **favour the exothermic reaction and shift to the left. [1 mark]**

e.   i. $K = \dfrac{[NO_2]^2}{[N_2O_4]}$ **[1 mark]**

       ii.   From the graph: $[NO_2] = 0.23$ M and $[N_2O_4] = 0.11$ M

$$K = \frac{[NO_2]^2}{[N_2O_4]} = \frac{(0.23)^2}{0.11} = \textbf{0.48 M [1 mark]}$$

Question 4 - [5 marks, 6 minutes]

a.   **[Mark allocation 1 mark for each correct answer: Total marks allocated 3 marks]**
   Increase the temperature of the reaction.
   Increase the concentration of the dilute nitric acid.
   Decrease the particle size of the limestone by grinding it to a finer state (**increase the surface area**).
   Increase the mass of limestone present.
   Add an appropriate catalyst. (This is not a common solution for this reaction.)

b.   **[Mark allocation 1 mark for each correct answer: Total marks allocated 2 marks]**
   Increase the temperature of the reaction. – Increasing the temperature results in **more particles having sufficient energy to overcome the activation barrier**, therefore a greater likelihood of a successful reaction when the particles collide.
   Increase the concentration of the dilute nitric acid. – Increasing the acid concentration results in more particles being present therefore a **higher chance of a reaction occurring**. (Care must be done with this as acids are corrosive.)
   Decrease the particle size of the limestone by grinding it to a finer state. – Decreasing the particle size for the limestone increases the surface area thereby **increasing the chance for a reaction to occur**.
   Increase the mass of limestone present. – Increasing the mass of limestone would result in more particles therefore **increase the chance for a reaction to occur**.
   Add an appropriate catalyst. – A catalyst will provide an alternative pathway for the reaction with a lower activation energy. Therefore **for a given temperature more particles will have sufficient energy to overcome this lower barrier** and hence the likelihood of a successful reaction increases.