Chemical Equilibrium

Chapter

7

1. **[7 marks]** *(2009:05)*

Solid magnesium hydroxide is added to a beaker of water. The water is stirred and the contents of the beaker left to settle. A saturated solution is formed, with undissolved magnesium hydroxide at the bottom of the beaker. The system can be shown by the following equation:

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2\ OH^-(aq)$$

(a) The system is allowed to come to equilibrium. Explain why the amount of solid present remains constant. *1.3* [1]

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(b) The changes indicated in the table below are now imposed on the system. Predict and explain the effect these changes have on the amount of solid magnesium hydroxide in the beaker once equilibrium is re-established. *1.9* [6]

Imposed change	**Effect on solid $Mg(OH)_2$ (write 'increase', 'decrease' or 'no change')**	**Explanation**
A little concentrated sodium hydroxide solution is added to the beaker		
Some sodium phosphate solution is added to the beaker		
More water is added to the beaker		

2. **[6 marks]** *(2010:26)*

Consider the following system:

$$CO(g) + 2\,H_2(g) \rightleftharpoons CH_3OH(g) \qquad \Delta H = -92\ kJ$$

(a) Predict whether the following changes will increase, decrease or have no effect on the rate of attainment of equilibrium. *1.1* [3]

Change	Effect
Decreasing the temperature	
Increasing the pressure of hydrogen	
Adding a catalyst	

(b) Predict whether the following changes will increase, decrease or have no effect on the equilibrium yield of the reaction. *1.9* [3]

Change	Effect
Increasing the temperature	
Increasing the pressure of the system	
Adding a catalyst	

3. **[2 marks]** *1.10 (2010:27)*

Write the equilibrium constant expression for the following equilbria:

(a) [1]

Equation	$BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$
Equilibrium constant expression	

(b) [1]

Equation	$2\,CrO_4^{2-}(aq) + 2\,H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + H_2O(\ell)$
Equilibrium constant expression	

4. **[4 marks]** *(2011:27b,c)*

Lactic acid produced by muscles during exercise, is found in many milk products and is used in the brewing of beer. It is also added to a number of canned food items as a buffer.

The equation for the reaction of lactic acid with water is shown below.

$\Delta H = -60\ \text{kJ mol}^{-1}$

The value of the equilibrium constant for the above reaction, at 25°C, is approximately 7.9×10^{-5}.

(a) State whether the ratio of organic products to organic reactants will be equal to one, less than one (< 1) or greater than one (> 1) for this system at equilibrium at 25°C. *1.11* [1]

(b) Predict the direction in which the equilibrium will shift immediately after the changes indicated in the table below. Write 'left', 'right' or 'no change'. *1.9* [3]

Change	Direction of initial equilibrium shift
Decreasing the temperature	
Adding hydrochloric acid	
Adding sodium hydroxide	

5. **[8 marks]** *1.9 (2011:28)*

Chloromethane can be produced industrially by the reaction of methanol and hydrogen chloride at high temperature in the presence of a catalyst. The equation for this reaction is shown below.

$$CH_3OH + HC\ell \rightleftharpoons CH_3C\ell + H_2O$$

The boiling points and melting points for each of the species involved in the reaction are shown below.

Species	Boiling point (°C)	Melting point (°C)
CH_3OH	65	-98
$HC\ell$	-85	-114
$CH_3C\ell$	-24	-98
H_2O	100	0

Write the phase, i.e., solid (*s*), liquid (*ℓ*) or gas (*g*), of each species in this system at the temperatures shown in the table below, and predict the effect of an increase in total pressure on this equilibrium at each of the temperatures.

Temperature (°C)	Phase (*s*, *ℓ* or *g*)				Shift in equilibrium (right, left or no change)
	CH_3OH	$HC\ell$	$CH_3C\ell$	H_2O	
-50					
40					
70					
110					

6. [6 marks] *1.9 (2012:29)*

The white solid bismuth oxychloride reacts with concentrated hydrochloric acid to establish the following equilibrium:

$$BiOC\ell(s) + 2\,H^+(aq) \rightleftharpoons Bi^{3+}(aq) + C\ell^-(aq) + H_2O(\ell)$$

Three test tubes of the equilibrium system, 'A', 'B' and 'C' were prepared by adding excess $BiOC\ell$ to concentrated hydrochloric acid.

Complete the table below by indicating the direction of the expected shift in equilibrium immediately following the changes stated in the table. Using Collision theory explain your choice of shift.

Test tube	Change	Direction of shift in equilibrium ('left', 'right' or 'no change')	Explanation
A	3 mL of water is added		
B	A few drops of concentrated nitric acid are added		
C	A few drops of concentrated silver nitrate solution are added		

7. **[6 marks]** *(2012:30)*

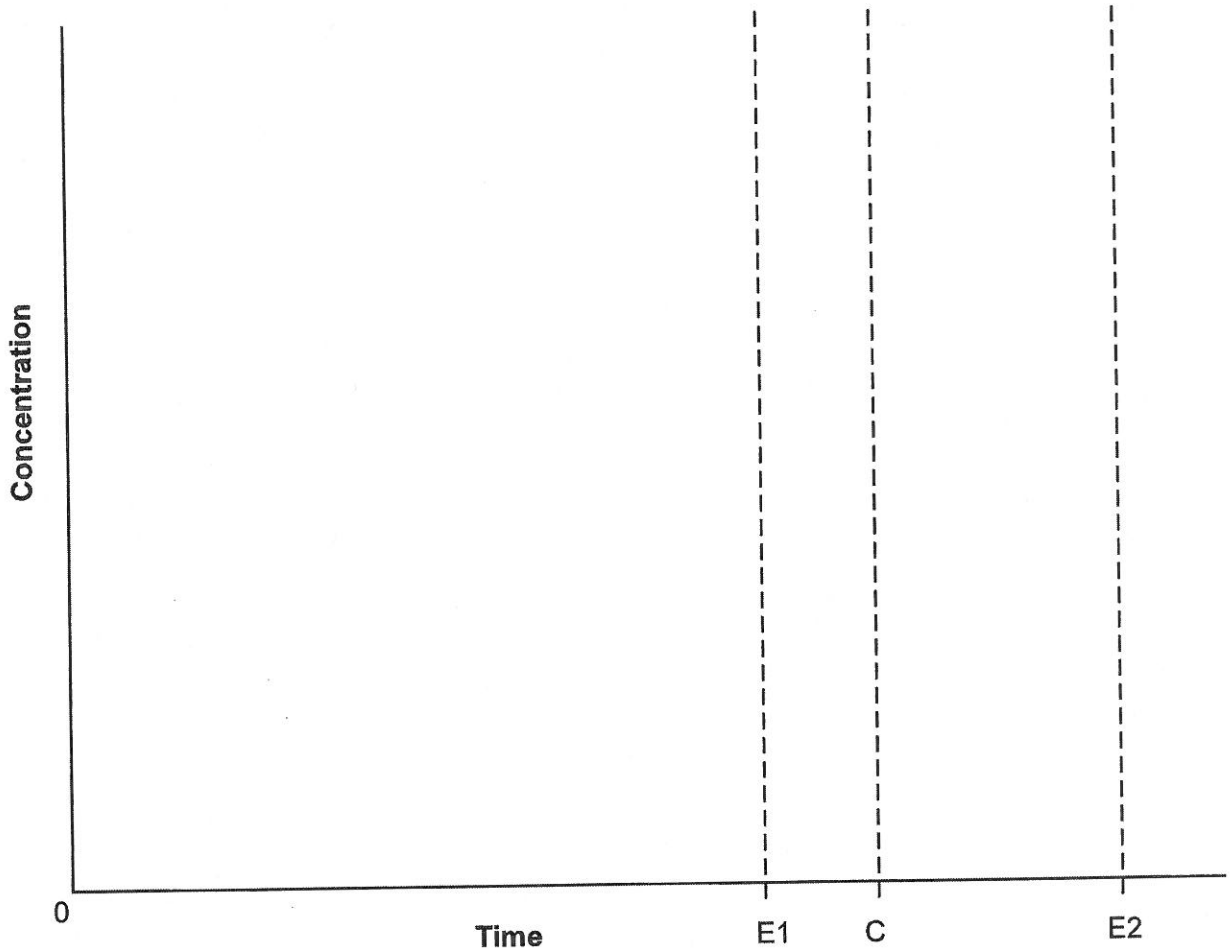
Silver chloride, $AgC\ell(s)$, is very sparingly soluble in water. However, it is soluble in ammonia solutions, due to the formation of the $[Ag(NH_3)_2]^+$ ion as shown in the equilibrium below:

$$AgC\ell(s) + 2\,NH_3(aq) \rightleftharpoons [Ag(NH_3)_2]^+(aq) + C\ell^-(aq)$$

The equilibrium constant, K, for this system is greater than 1 (>1).

A student mixes the reactants at time t = 0.

(a) On the axes below, draw separate curves to show how the concentrations of $NH_3(aq)$ and $[Ag(NH_3)_2]^+(aq)$ change with time as the system approaches, and finally reaches, equilibrium (Time E1). Clearly label your curve for $NH_3(aq)$ and your curve for $[Ag(NH_3)_2]^+(aq)$. Continue your curves from Time E1 to Time C. *1.4* [3]



(b) At Time = C, as shown on the axis, a small quantity of concentrated $NaC\ell$ solution is added to the system, and the system is then again allowed to reach equilibrium at Time E2. On the same axes above, show how the concentrations of $NH_3(aq)$ and $[Ag(NH_3)_2]^+(aq)$ would change in response to the addition of $NaC\ell$ solution from Time C until equilibrium is reached at Time E2. *1.9* [3]

8. **[4 marks]** *1.10 (2013:29)*

Write the equation and the expression for the equilibrium constant for each of the equilibrium processes below.

Equilibrium process	**Equation**	**Equilibrium constant expression**
Vaporisation of water		
Dissolution of solid aluminium sulfate in water		

9. **[8 marks]** *1.8 (2013:30)*

Consider the following system at equilibrium.

$$4\ NH_3(g) + 5\ O_2(g) \rightleftharpoons 4\ NO(g) + 6\ H_2O(g) + 920\ kJ$$

Indicate in the table below whether there would be an increase, decrease, or no change in the concentration of $NH_3(g)$ after the changes given in the table are imposed on the system and **equilibrium has been re-established**. Provide a brief explanation for the observation.

Change	Change in concentration of $NH_3(g)$ (circle the correct response)	Brief explanation
The volume of the reaction vessel is doubled	• increase • decrease • no change	
The temperature of the reaction system is doubled	• increase • decrease • no change	
$N_2(g)$ is injected into the reaction system while keeping the volume constant	• increase • decrease • no change	
Water vapour is injected into the reaction system while keeping the volume constant	• increase • decrease • no change	

10. **[6 marks]** *1.4, 1.7 (2015:30)*

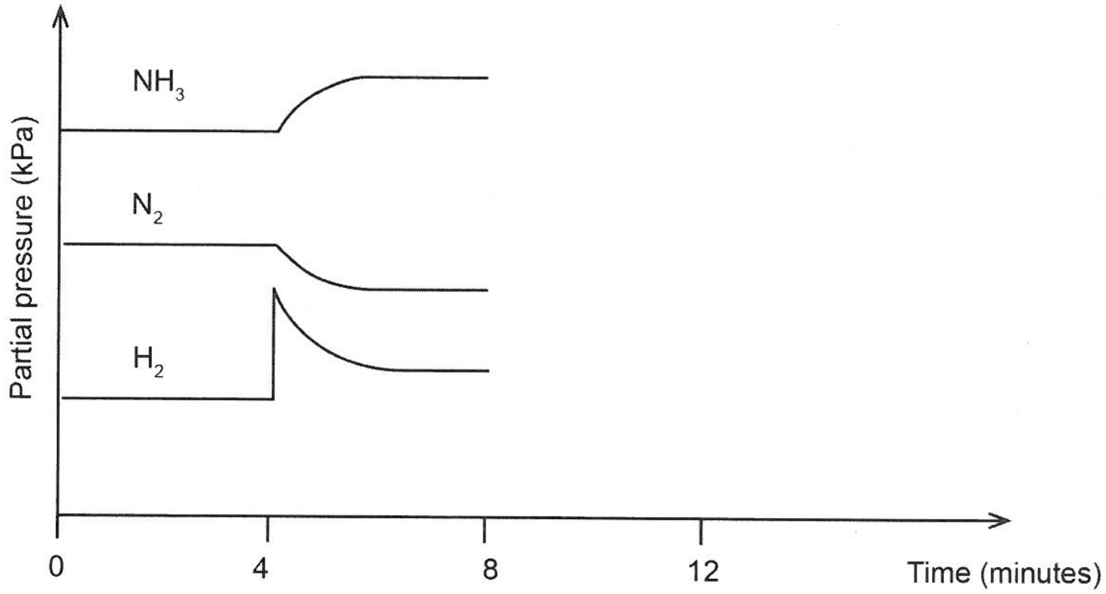
Ammonia exists in equilibrium with hydrogen and nitrogen as shown by the following exothermic equation.

$$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g) \qquad \Delta H = -\,92 \text{ kJ mol}^{-1}$$

As they exist in the gaseous state, the relative concentrations can be given in terms of the partial pressure (kPa) of each gas.

Nitrogen, hydrogen and ammonia gases are placed in a rigid container and allowed to reach equilibrium. The graph below shows the partial pressures of the gaseous system initially at equilibrium. After the experiment operates for 4 minutes, a change is imposed upon it.

Partial pressures of NH_3, N_2 and H_2 over time



(a) What characteristic of equilibrium is indicated on the graph by the section from 0 to 4 minutes? [1]

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(b) A change was imposed on the system at the 4 minute mark. What imposed change could have produced the results indicated on the graph? [1]

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(c) The system was **suddenly** cooled at 8 minutes and then reached equilibrium again at 12 minutes. Using this information, complete the graph above from the 8 to the 12 minute mark. [4]

11. [16 marks] *(2015:38)*

The two different coloured cobalt(II) complex ions, $Co(H_2O)_6^{2+}$ and $CoC\ell_4^{2-}$, exist together in equilibrium in solution in the presence of chloride ions. This is represented by the equation below.

$$\underset{\text{pink}}{Co(H_2O)_6^{2+}(aq)} + 4\ C\ell^-(aq) \rightleftharpoons \underset{\text{blue}}{CoC\ell_4^{2-}(aq)} + 6\ H_2O(\ell)$$

An experiment is conducted to investigate the effects on the equilibrium position by imposing a series of changes on the system. The shift in equilibrium position can be indicated by any colour change of the solution.

Colour chart	
Species	**Colour**
$Co(H_2O)_6^{2+}(aq)$	pink
$CoC\ell_4^{2-}$ (aq)	blue
Initial equilibrium mixture	purple

After a 3.00 mL sample of an initial equilibrium mixture was placed in each of three test tubes, changes to each system were made by adding a different substance, as indicated in the table below.

Test tube	Substance added to the test tube
1	10 to 12 drops of distilled water
2	20 to 25 drops of concentrated hydrochloric acid
3	20 to 25 drops of 0.200 mol L^{-1} silver nitrate solution, $AgNO_3(aq)$

(a) Complete the table below by predicting the:
- change in concentration, if any, of each of the ions in solution compared to the initial solution, after a new equilibrium position is reached.
- colour change, if any, that takes place from the initial purple-coloured solution. [6]

Additions to the test tube	Change in concentration from initial equilibrium to final equilibrium (increase, decrease, unchanged)			Colour favoured (pink, blue or unchanged)
	$[Co(H_2O)_6^{2+}]$	$[C\ell^-]$	$[CoC\ell_4^{2-}]$	
1. add $H_2O(\ell)$				
2. add $HC\ell(aq)$				
3. add $AgNO_3(aq)$				

(b) Other than a colour change, what else should be observed in test tube 3? [1]

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(c) Using Collision Theory, explain your predicted observations when hydrochloric acid is added to test tube 2 [3]

Another experiment was conducted to investigate the effect that changing the temperature had on the equilibrium mixture. When 3.00 mL of the original equilibrium mixture was placed in a test tube and then in an ice bath, the solution became pink.

(d) Determine whether the forward reaction, as illustrated by the equation below, is exothermic or endothermic. Use Le Châtelier's Principle to justify your answer. [4]

$$\underset{\text{pink}}{Co(H_2O)_6^{2+}(aq)} + 4\,C\ell^-(aq) \rightleftharpoons \underset{\text{blue}}{CoC\ell_4^{2-}(aq)} + 6\,H_2O(\ell)$$

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(e) State **one** specific hazard to the environment that the disposal of chemicals from this experiment poses and state what could be done in the laboratory to reduce this hazard. [2]

12. **[6 marks]** *1.0 (2016 SP:34)*

Ocean acidification results from carbon dioxide dissolving in water and an equilibrium being established between the water and carbon dioxide to produce carbonic acid, (H_2CO_3).

(a) Write a balanced equation for this equilibrium. [2]

(b) The formation of carbonic acid leads to an increase in the hydronium ion (H_3O^+) concentration in water. Show the equilibrium that results in the formation of hydronium ions when carbonic acid reacts with water. [1]

(c) State **one** problem ocean acidification is causing for marine organisms. Explain how this problem arises and support your answer with an appropriate balanced equation. [3]

13. **[18 marks]** *1.11 (2016:41)*

Nitrogen dioxide is toxic to humans when inhaled and is a significant component of air pollution. It can be formed by the combustion of nitrogen in the air at high temperatures; firstly forming nitric oxide $NO(g)$ and on further oxidation, forming nitrogen dioxide, $NO_2(g)$. The overall equation for this process is given here:

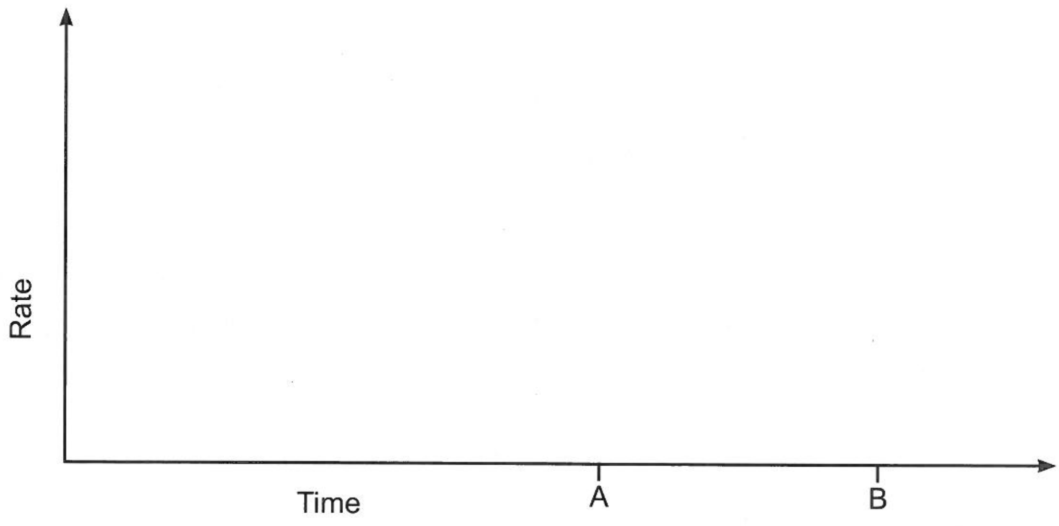
$$N_2(g) + 2\,O_2(g) \rightleftharpoons 2\,NO_2(g)$$

The following questions relate to the equilibrium system represented by this equation.

(a) Write the equilibrium expression for this reaction when it is in equilibrium. [2]

(b) Assuming all other conditions remain constant, what happens to the equilibrium constant after the pressure of the system is lowered and equilibrium is re-established? [1]

(c) (i) On the axes below, draw the forward (—) and reverse (- - -) reaction rates, starting at the moment the oxygen and nitrogen gases begin to react with each other until after equilibrium has been established at time A. Continue the graph until time B. [3]



(ii) On the same axes above, draw and label clearly the effect of conducting the same reaction at a higher temperature. [2]

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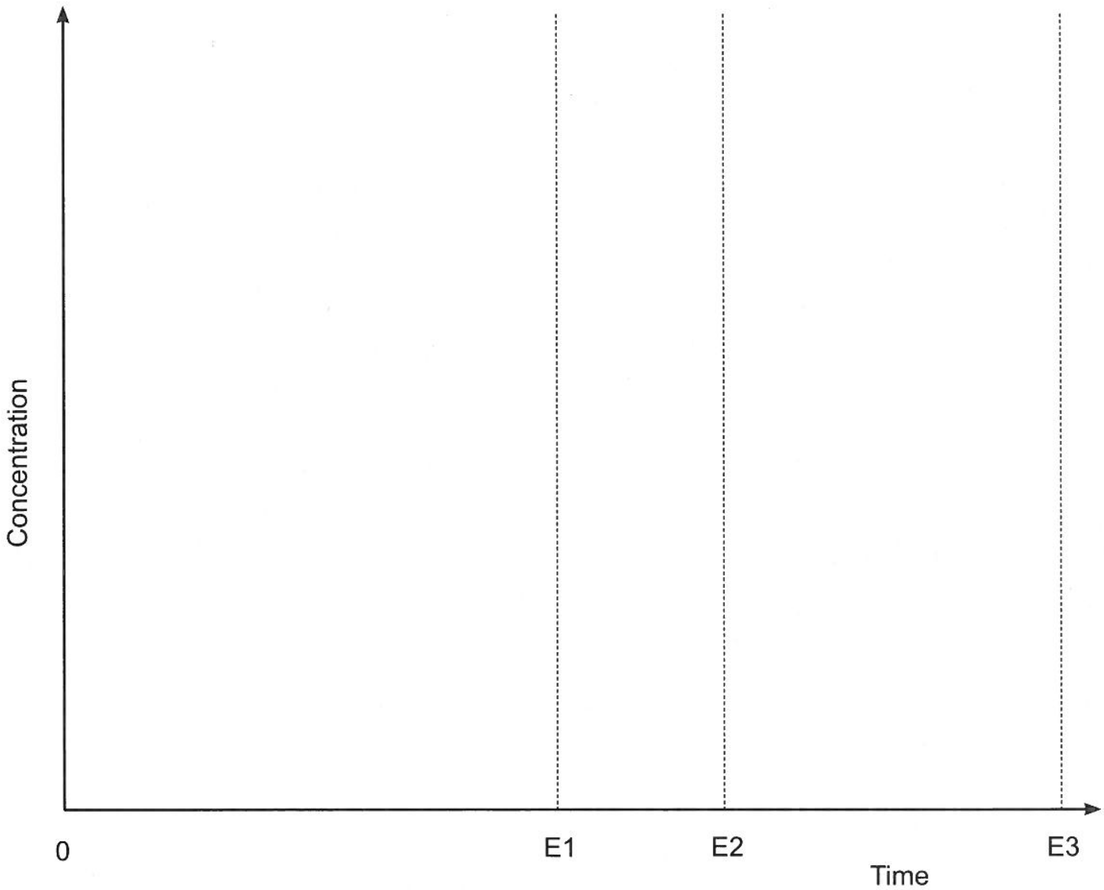
(d) On the axes below, draw separate curves to show how the concentrations of the **three** gases change with time, starting at the moment the oxygen and nitrogen gases begin to react with each other until the system reaches equilibrium at Time E1. Continue the graph from Time E1 to Time E2. Assume that the initial concentrations of oxygen and nitrogen are identical.

Label clearly the line for each gas. [5]



(e) At Time E2 shown on the axis, the reaction vessel is doubled in volume, and the system is then again allowed to reach equilibrium at Time E3. On the same graph above, show how the concentrations of the three gases would change in response to the change in volume, from Time E2 until equilibrium is re-established at Time E3. [3]

The reaction between nitrogen gas and oxygen gas occurs at high temperatures such as those found in the combustion engines of cars. The atmosphere is composed of 78% nitrogen and 21% oxygen and has been stable for millions of years.

(f) What does the stability of this composition indicate about the equilibrium constant and energy requirements of the reaction between nitrogen and oxygen gases? [2]

Chapter 7: Chemical Equilibrium

1.(2009:05) a) The magnesium hydroxide solution is saturated. It has come to equilibrium. This means there are no further changes in the concentrations since the rate of dissolution is equal to the rate of re-crystallisation, allowing no net change in the amount of the undissolved solid present.

b) $Mg(OH)_2(s) \rightarrow Mg^{2+}(aq) + 2\,OH^-(aq)$

Note sections of syllabus 1.7 to 1.9 make it clear that you cannot use Le Châtelier's Principle to 'explain' only to 'predict'.

Imposed change	**Effect on solid $Mg(OH)_2$ (Write 'increase', 'decrease' or 'no change')**	**Explanation**
A little concentrated sodium hydroxide solution is added	The amount of solid increases	Increasing the $[OH^-](aq)$ will increase the number of collisions per second on the right hand side and increase the rate of the reverse reaction. Thus the equilibrium will move to the LHS.
Some sodium phosphate solution is added to the beaker	The amount of solid decreases	Adding $PO_4^{3-}(aq)$ ions will precipitate out some of the $Mg^{2+}(aq)$ ions. Thus there will be less concentration of $Mg^{2+}(aq)$ and less collisions per second. So there will be a lowered rate of reverse reaction. The forward rate is unchanged so the equilibrium will shift to the RHS and there will be less $Mg(OH)_2(s)$ present.
More water is added to the beaker	The amount of solid decreases	By adding water the solution is diluted. The concentration of $Mg^{2+}(aq)$ and $OH^-(aq)$ is lowered. Thus there will be less collisions per second and the rate of the reverse reaction will drop. The forward reaction rate will be unaffected so more $Mg(OH)_2$ will dissolve.

2.(2010:26)

a) $CO(g) + 2\,H_2(g) \rightleftharpoons CH_3OH(g)$

Change	**Effect on the rate**
Decreasing the temperature Increasing the pressure of hydrogen Adding a catalyst	Decrease Increase Increase

b)

Change	**Effect on the yield**
Increasing the temperature Increasing the pressure of the system Adding a catalyst	Decrease Increase No change

3.(2010:27)

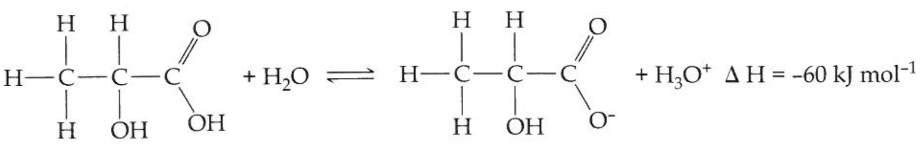
a) Equation: $BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$

Equilibrium expression: Keq. = $[Ba^{2+}(aq)] \times [SO_4^{2-}(aq)]$

b) Equation: $2\,CrO_4^{2-}(aq) + 2\,H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + H_2O(l)$

Equilibrium expression: Keq. = $\{[Cr_2O_7^{2-}(aq)] \div [CrO_4^{2-}(aq)]^2 \times [H^+(aq)]^2\}$

The equation for the reaction of lactic acid with water is given here.



4.(2011:27b,c) a) The value for the equilibrium constant for the above reaction at 25°C, is approximately 7.9×10^{-5}. Since the equilibrium constant is far less than one, the ratio of the organic products to organic reactants is less than one.
b) The direction in which the equilibrium will shift immediately after the changes indicated in the table:

Change	Direction of initial equilibrium shift
Decreasing the temperature	To the right. (As this is an exothermic reaction, a decrease in temperature increases the yield while decreasing the rate).
Adding hydrochloric acid	To the left. (Adding H^+ ions increases the concentration of one of the products which shifts the equilibrium to the left decreasing the yield of the reaction as written).
Adding sodium hydroxide	To the right. (The added H^+ ions react with H_3O^+ ions and decrease its concentration. Hence the equilibrium shifts to the right to increase the yield of the reaction as written).

5.(2011:28)

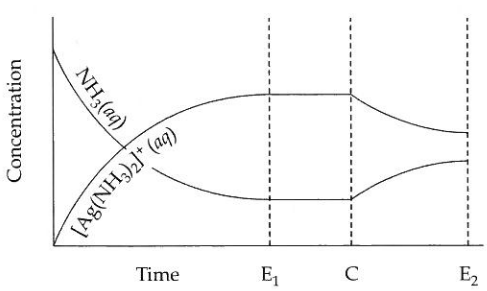
Temperature (°C)	Phase				Shift in equilibrium (right, left or no change)
	CH_3OH	$HC\ell$	$CH_3C\ell$	H_2O	
-50	L	G	L	S	Right
40	L	G	G	L	No change
70	G	G	G	L	Right
110	G	G	G	G	No change

6.(2012:29) The white solid in this reaction is in equilibrium with concentrated $HC\ell$. The reaction is:
$BiOC\ell(s) + 2\,H^+(aq) \rightleftharpoons Bi^{2+}(aq) + C\ell^-(aq) + H_2O(\ell)$
Three test tubes of equilibrium solutions A, B and C were prepared by adding excess $BiOC\ell$ to concentrated H_2SO_4. When changes are made as shown in column two, completion of the table is required.

Test tube	Change	Direction of shift in equilibrium ('left' 'right' or 'no change')	Explanation
A	3 mL of water is added	No change	The water lowers the concentration of all species. As there are equal moles on each side both sides are equally affected and both the forward and reverse rates are lowered equally.
B	A few drops of concentrated nitric acid are added	Shifts to the right	Adding concentrated nitric acid does not dilute the solution as in (A) but increases the concentration of H^+ ions. Thus the forward reaction rate increases while the backward is unchanged. So the equilibrium shifts to the right.
C	A few drops of concentrated silver nitrate are added	Shifts to the right	Added Ag^+ ions react with $C\ell^-$ ions to produce $AgC\ell(s)$ precipitate. Therefore the concentration of $C\ell^-$ falls and the reverse rate of reaction falls. The forward reaction is unchanged and thus the equilibrium shifts to the right resulting in more $BiOC\ell$ dissolving

7.(2012:30) $AgC\ell$ solid dissolves in ammonia solution to form a complex ion as shown in the following equilibrium reaction: $AgC\ell(s) + 2\,NH_3(aq) \rightleftharpoons [Ag(NH_3)_2]^+(aq) + C\ell^-(aq)$. The equilibrium constant (K_c) is greater than 1. This means that the concentration of the products is greater than the concentration of the reactants.
a) The following graph shows how the concentrations change as the system approaches equilibrium at time E, and continues on to time C. The graph also shows the concentration changes to NH_3 and $[Ag(NH_3)_2]^+(aq)$ when $NaC\ell$ is added to the system.

$AgC\ell(s) + 2\,NH_3(aq) \rightleftharpoons [Ag(NH_3)_2]^+\,(aq) + C\ell^-(aq)$



b) As $NaC\ell$ is added to the system, the chloride ion concentration increases and the equilibrium shifts to the left, decreasing the $[Ag(NH_3)_2]^+$ and increasing the concentration of $[NH_3]$. More $AgC\ell$ is precipitated.

8.(2013:29)

Equilibrium Process	Equation	Equilibrium Constant Expression
Vaporisation of water	$H_2O(\ell) \rightleftharpoons H_2O(g)$	$K = [H_2O(g)]$
Dissolution of solid aluminium sulfate in water	$A\ell(SO_4)_3\ (s) \rightleftharpoons$ $2\ A\ell^{3+}(aq) + 3\ SO_4^{2-}(aq)$	$K = [A\ell^{3+}(aq)]^2\ [SO_4^{2-}(aq)]^3$

9.(2013:30) Note sections of syllabus 1.7 to 1.9 make it clear that you cannot use Le Châtelier's Principle to 'explain' only to 'predict'.

The reaction should be written like this.

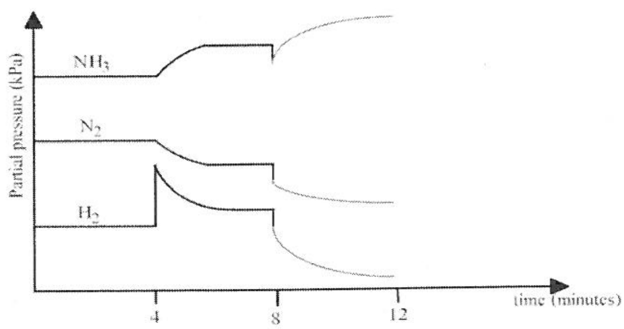
$$4\ NH_3(g) + 5\ O_2(g) \rightleftharpoons 4\ NO(g) + 6\ H_2O(g) + \text{Heat}$$

Change	Change in concentration of $NH_2(g)$	Brief explanation
The volume of the reaction vessel is doubled	Decrease	The pressure in the system decreases. Therefore the volume of the system increases. As c = n/V the concentrations decrease for all species. So both the forward and reverse reactions will have a lower rate. The reverse more so because there are 10 mol on the RHS. Therefore there will be more product formed.
The temperature of the reaction system is doubled	Increase	The increase in temperature of the system favours the endothermic reaction. In an energy diagram for this reaction the E_a of the reverse reaction is larger than the E_a of the forward reaction. At an elevated temperature more product particles can achieve E_a and the reverse rate will rise. Thus the reaction moves to the LHS.
$N_2(g)$ is injected into the reaction system while keeping the volume constant	No change	You will often see questions with an unusual substance added. Sometimes it reacts (sometimes by precipitation) but in this case N_2 gas does not participate in any reaction. It increases the pressure but does not change the volume. So as c = n/V there is no change in the concentration and no change in the rate of reaction. No change in NH_3 concentration.
Water vapour is injected into the reaction system while keeping the volume constant	Increase	Usually water is in the (ℓ) liquid state and does not take part in the equilibrium considerations. Here, however, it is in the vapour state, there is an increase in the concentration of product. This increases the rate of the reverse reaction increasing the concentration of ammonia. The equilibrium moves to the LHS.

10.(2015:30) a) Between 0 and 4 minutes, the partial pressures of all gases were constant.

b) H_2 gas was added.

c) There was an immediate decrease, then the $[NH_3]$ rose and the others fell in stoichiometric ratios, then equilibrium was established.



11.(2015:38) a)

Additions to the test tube	Change in concentration from initial equilibrium to final equilibrium (increase, decrease, unchanged)			Colour favoured (pink, blue or unchanged)
	$[Co(H_2O)_6^{2+}]$	$[C\ell^-]$	$[CoC\ell_4^{2-}]$	
1. add $H_2O(\ell)$	decrease	decrease	decrease	pink
2. add $HC\ell\ (aq)$	decrease	increase	increase	blue
3. add $AgNO_3(aq)$	increase	decrease	decrease	pink

b) A white precipitate forms

c) Adding $HC\ell(aq)$ increases the concentration of chloride ions and the number of collisions between $Co(H_2O)_6^{2+}$ and $C\ell^-$ increases. The rate of the forward reaction increases relative to the reverse reaction and hence equilibrium shifts to the right. This leads to a greater concentration of the blue $CoC\ell_4^{2-}$ and a lower concentration of the pink $Co(H_2O)_6^{2+}$ ion, so the solution looks more blue.

d) Decreasing the temperature shifts the equilibrium position to the left favouring the production of $Co(H_2O)_6^{2+}(aq)$, a pink solution. According to Le Châtelier's Principle decreasing the temperature favours the exothermic reaction to oppose the change. Since the reverse reaction has been favoured, that is the exothermic reaction. So the forward reaction is endothermic.

e) Any hazard relevant to this specific experiment. eg the disposal of concentrated hydrochloric acid. Any method that is relevant eg dilute with large amounts of water when emptying down the sink.

12.(2016 SP:34) Calcium carbonate does dissolve at the rate of 0.13 g/L and the ocean is a saturated solution of calcium carbonate (equation 3) and it is this source of carbonate ions that corals and shellfish use to create their exoskeleton.

Equation 1 $\quad CaCO_3\ (s) \rightleftharpoons Ca^{2+}(aq) + CO_3^{2-}(aq) \quad K = 3.3 \times 10^{-9} = \dfrac{[Ca^{2+}]\ [CO_3^{2-}]}{1}$

Atmospheric carbon dioxide is in equilibrium with dissolved carbon dioxide in the ocean

Equation 2 $\quad CO_2(g) \rightleftharpoons CO_2(aq)$

The dissolved carbon dioxide undergoes hydrolysis (equation 2).

Equation 3 $\quad CO_2(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq) \quad K_1 = \dfrac{[H_3O^+]\ [HCO_3^-]}{[CO_2]}$

As the concentration of dissolved CO_2 increases the pH of the ocean falls, becoming more acidic.

The excess H^+ formed as the ocean becomes more acidic (equation 3) reacts with soluble carbonate in the ocean (equation 4) resulting in an increase in the concentration of HCO_3^- (equation 4).

Equation 4 $\quad CO_3^{2-}(aq) + H_3O^+(\ell) \rightleftharpoons HCO_3^-(aq) \quad K_2 = \dfrac{[H_3O^+]\ [CO_3^{2-}]}{[HCO_3]}$

Coral body is built from calcium carbonate when the coral combine with carbonate ions from the water.

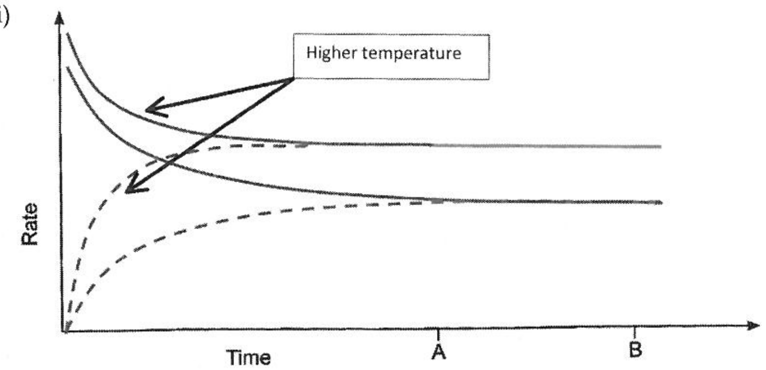
Equation 5 $\quad Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

In accordance with Le Châtelier's Principle as the hydrogen ion removes carbonate ion in equation 4 it will force equation 1 to the right to replace some of the carbonate. Not all is replaced because the $[Ca^{2+}]$ is rising and to maintain K the carbonate decreases. There is an effective lowering of carbonate ion concentration.

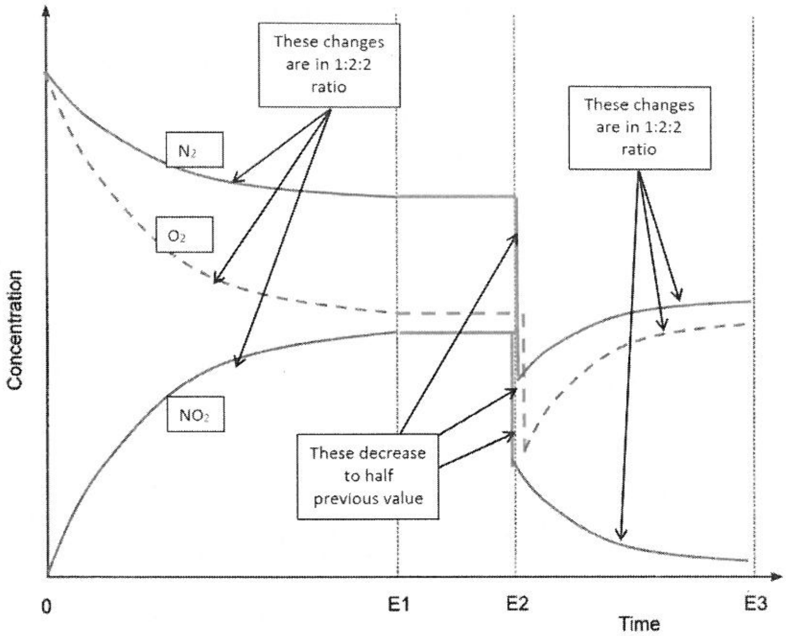
13.(2016:41) a) $K = \dfrac{pNO_2(g)^2}{pN_2(g)pO_2(g)^2}$ or $K = \dfrac{[NO_2]^2}{[N_2][O_2]^2}$

b) No change in K

c) i) and ii)



d) and e)



f) As the reactants are favoured, the equilibrium constant would be small. The activation energy would be large.

Chapter 8: Acids and Bases

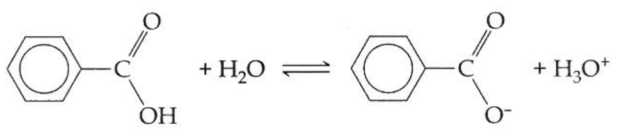
1.(2010:28)

a) $NH_3 + NH_3 \rightleftharpoons NH_4^+ + NH_2^-$

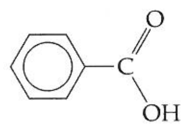
Base, Acid, conjugate acid, conjugate base

b) Kc. = 1×10^{-30}. This is a very low 'K' value which indicates that the reaction does not produce any significant amount of products. It is assumed that the temperature at which this equilibrium constant is calculated is more than 0°C. The reaction is also an endothermic process. An increase in temperature will favour the forward reaction and increase the concentration of the products for an endothermic reaction. This will effectively increase the Kc. value above 1×10^{-30}.

2.(2010:29) a) $C_6H_5OH + H_2O \rightleftharpoons C_6H_5COO^- + H_3O^+$



b) In the acidic environment of the stomach which produces gastric juice that contains hydrochloric acid, the concentration of H_3O^+ ions is rather high. This shifts the equilibrium to the left favouring more reactants. Furthermore, benzoic acid is an organic acid and the extent of its ionisation is much less than inorganic acids. Therefore, the predominant species in the acidic environment of the stomach would be benzoic acid. The structure of benzoic acid is given below.



c) Benzoic acid and benzoate ion can exist as a buffer system as they are a weak acid-conjugate base pair. It can resist a change in pH by reacting with both OH^- ions and H_3O^+ ions as shown below, resulting in equilibrium shifts either way.

Reaction with H_3O^+ ions: $H_3O^+ + C_6H_5COO^- \rightleftharpoons H_2O + C_6H_5COOH$

Reaction with OH^- ions: $OH^- + C_6H_5COOH \rightleftharpoons H_2O + C_6H_5COO^-$

The equilibrium reaction that occurs both ways is shown by a single equation below.