Questions

Module 5: Equilibrium and Acid Reactions

Multiple-choice questions: 1 mark each

1. The following equilibrium is set up in a sealed reaction vessel.

$$N_2O_4(g) \rightleftharpoons 2NO_2(g) \qquad \Delta H = +54.8 \text{ kJ mol}^{-1}$$

Which of the following would INCREASE the yield of nitrogen dioxide?

(A) Adding a catalyst to the reaction vessel

(B) Decreasing the volume of the reaction vessel

(C) Raising the temperature of the reaction vessel

(D) Increasing the pressure by adding argon to the reaction vessel

2013 HSC Q10

2. When chloride ions are added to a solution containing $Co(H_2O)_6^{2+}(aq)$, the following equilibrium is established.

$$\underset{\text{Pink}}{Co(H_2O)_6^{2+}(aq)} + 4Cl^-(aq) \rightleftharpoons \underset{\text{Blue}}{CoCl_4^{2-}(aq)} + 6H_2O(\ell)$$

Which of the following statements about the colour of the solution is true?

(A) Diluting the solution with water will make it turn blue.

(B) If the reaction is exothermic, heating the solution will make it turn blue.

(C) If the reaction is endothermic, cooling the solution will make it turn pink.

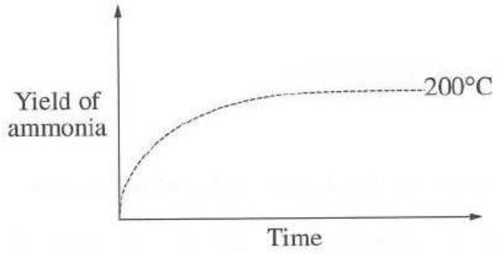
(D) Adding a large amount of solid potassium chloride to the solution will make it turn pink.

2011 HSC Q13

3. Ammonia is produced from hydrogen and nitrogen, according to the equation:

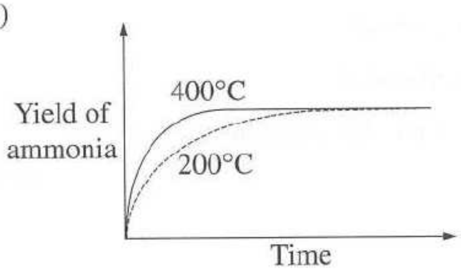
$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H \text{ is negative}$$

The graph shows the yield of ammonia produced at 200°C and 100 kPa.



Which graph shows a correct comparison of the yield of ammonia produced at a temperature of 400°C and 100 kPa with the yield produced at 200°C and 100 kPa?

(A)



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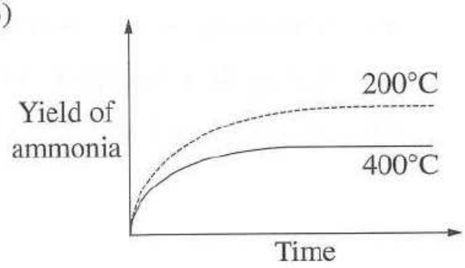

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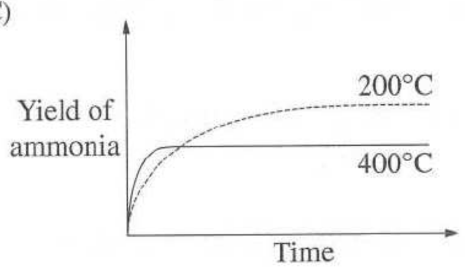

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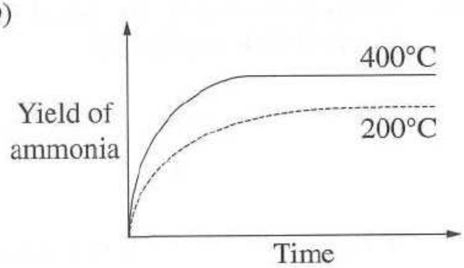


2002 HSC Q14

4. In which of the following reactions is the amount of products increased by an increase in pressure?

(A) $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$

(B) $Cl_2(g) + 3F_2(g) \rightleftharpoons 2ClF_3(g)$

(C) $2O_3(g) \rightleftharpoons 3O_2(g)$

(D) $H_2CO_3(aq) \rightleftharpoons H_2O(l) + CO_2(g)$

5. Phosgene is prepared from the reaction of carbon monoxide and chlorine in the presence of a catalyst:

$$CO(g) + Cl_2(g) \rightleftharpoons COCl_2(g) \qquad \Delta H = -9.93 \text{ kJ mol}^{-1}$$

Which of the following sets of conditions would produce the highest yield of phosgene?

(A) High temperature, high pressure

(B) Low temperature, low pressure

(C) Low temperature, high pressure

(D) High temperature, low pressure

2004 HSC Q10

6. Which of the following changes will always shift this equilibrium reaction to the right?

$$2HI(g) \rightleftharpoons H_2(g) + I_2(g) \qquad \Delta H = -52 \text{ kJ}$$

(A) Adding a catalyst

(B) Increasing the pressure

(C) Increasing the temperature

(D) Adding more of the reactant

2008 HSC Q7

7. An understanding of Le Chatelier's principle is important in the chemical industry. Which prediction can be made using this principle?

(A) The identity of products of a chemical reaction

(B) The effect of changes in temperature on the rates of reactions

(C) The effect of catalysts on the position of equilibrium reactions

(D) The effect of changes in the concentration of chemical substances in equilibrium

2001 HSC Q9

8. A chemical reaction is considered to be a 'system', while everything outside of the system is considered to be its 'surroundings'. Such systems can be open or closed.

Which of these statements correctly applies to an equilibrium system?

(A) A system that can only exchange energy with its surroundings.

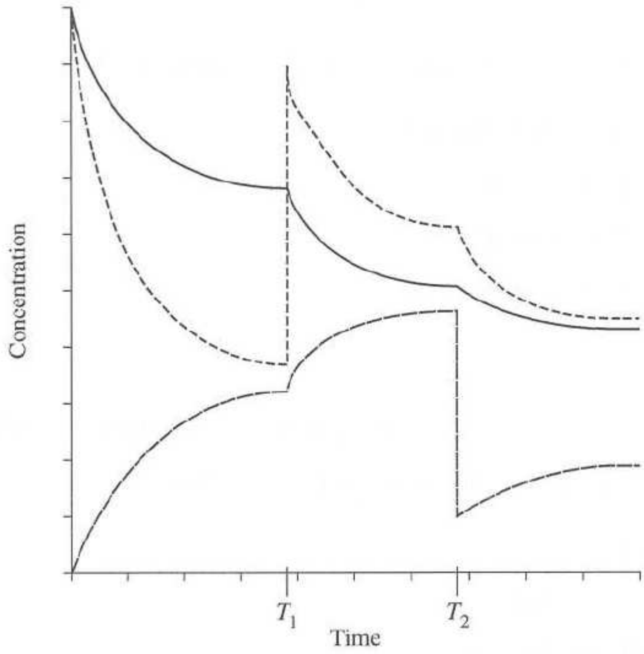
(B) A system that only occur in the forward direction.

(C) A system that can exchange reactants, products and energy with its surroundings.

(D) A system that can occur in the reverse direction.

9. The graph shows the concentrations over time for the system:

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



What has happened at times T_1 and T_2?

	T_1	T_2
(A)	H_2 added	CH_3OH removed
(B)	CO added	CH_3OH removed
(C)	H_2 added	CO removed
(D)	CO added	CO and H_2 removed

2012 HSC Q16

10. Consider the following reaction at equilibrium.

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

What would be the effect of a decrease in pressure on this system?

(A) Heat will be absorbed.

(B) The equilibrium will not be disturbed.

(C) The concentration of NH_3 will increase.

(D) The reverse rate of reaction will decrease.

2007 HSC Q13

11. Chromate and dichromate ions form an equilibrium according to the following equation.

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

Which solution would increase the concentration of the chromate ion $\left(CrO_4^{2-}\right)$ when added to the equilibrium mixture?

(A) Sodium nitrate

(B) Sodium chloride

(C) Sodium acetate

(D) Ammonium chloride

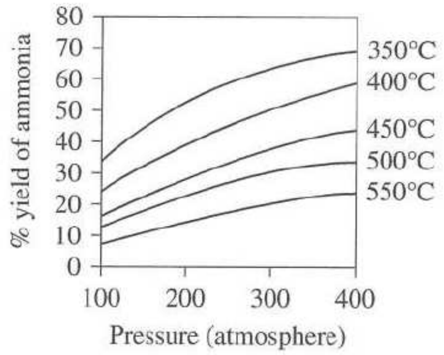
2010 HSC Q18

12. This Haber process is shown by the equation:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H \text{ is negative}$$

Which graph shows how pressure and temperature affect the yield of ammonia produced by the Haber process?

(A)



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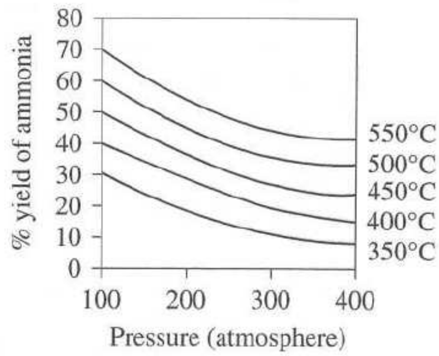

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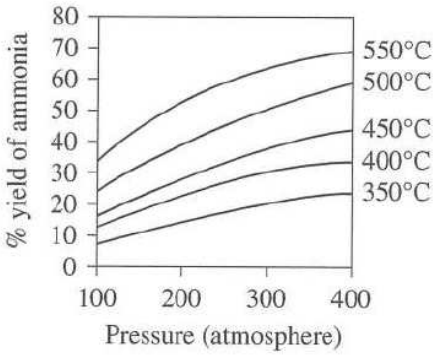

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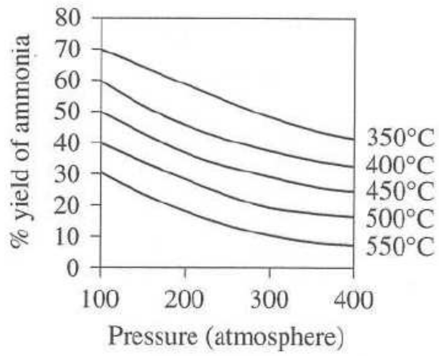
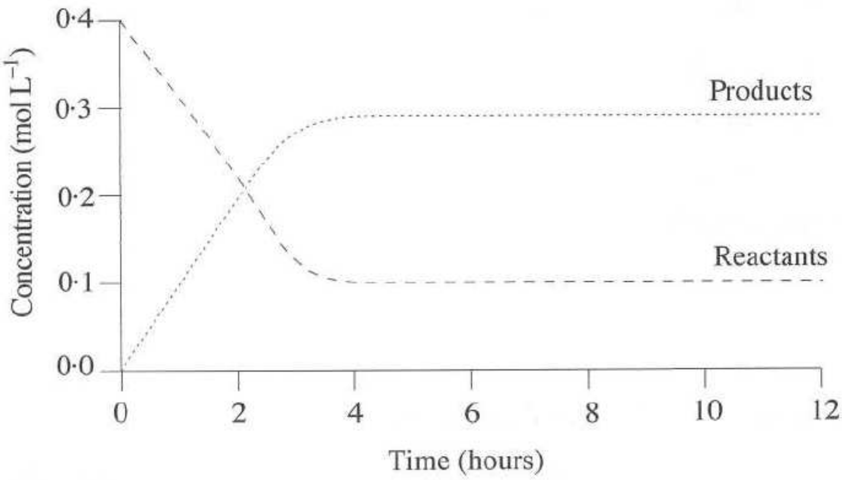


2006 HSC Q15

13. The graph below shows how the concentration of reactants and products change over time for the reaction:

$$CH_3COOC_2H_5 + H_2O \rightleftharpoons CH_3COOH + C_2H_5OH$$



From this graph we can determine that

(A) the equilibrium concentrations were 0·2 mol L^{-1}.

(B) the forward reaction stopped after four hours.

(C) the system reached equilibrium after two hours.

(D) the reaction did not go to completion.

1996 HSC Q6

14. The correct expression for the equilibrium constant K for the equation

$$2Cl_2(g) + 2H_2O(g) \rightleftharpoons 4HCl(g) + O_2(g)$$

is

(A) $\dfrac{[H_2O]\,[Cl_2]}{[HCl]\,[O_2]}$

(B) $\dfrac{[HCl]\,[O_2]}{[H_2O]\,[Cl_2]}$

(C) $\dfrac{2[H_2O]\,2[Cl_2]}{4[HCl]\,[O_2]}$

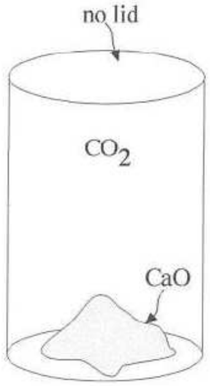
(D) $\dfrac{[HCl]^4\,[O_2]}{[H_2O]^2\,[Cl_2]^2}$

1995 HSC Q6

15. The conversion of calcium carbonate to calcium oxide and carbon dioxide is a reversible reaction and will reach equilibrium under certain conditions.

Which of the following diagrams shows that the system may have reached equilibrium?

(A)



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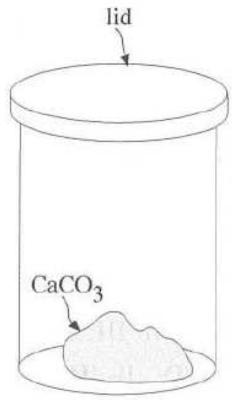

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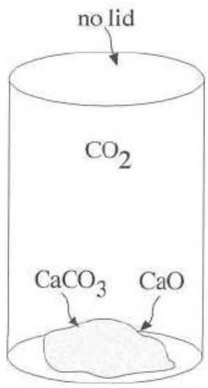

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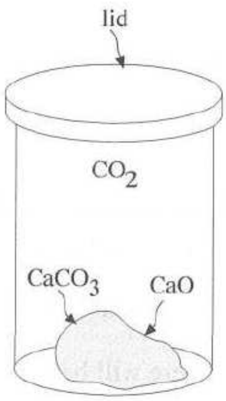


1995 HSC Q11

16. In a chemical equilibrium system, a chemical reaction can be either open or closed, and the changes taking place can either be reversible or non-reversible.

Which of the following statements is incorrect?

(A) A non-reversible reaction allows the reactants and products to go back and forth in both directions.

(B) An open system interacts with its environment and so energy and matter can move between it and its surroundings.

(C) In a reversible reaction, the products once they have been formed, can react again and so re-form the reactants.

(D) A closed system does not interact with its environment and so can only exchange energy with its surroundings.

17. Arrange the following reactions in order of their increasing tendency to reach completion.

I	$4NH_3(g) + 3O_2(g) \rightleftharpoons 2N_2(g) + 6H_2O(g)$	$K = 1{\cdot}0 \times 10^{228}$
II	$2HF(g) \rightleftharpoons H_2(g) + F_2(g)$	$K = 1{\cdot}0 \times 10^{-13}$
III	$2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)$	$K = 4{\cdot}7 \times 10^{-4}$
IV	$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$	$K = 5{\cdot}0 \times 10^{-31}$

(NOTE: For each reaction, the equilibrium constant was determined under different conditions.)

(A) I, III, II, IV
(B) III, II, IV, I
(C) IV, II, III, I
(D) I, IV, II, III

1996 HSC Q9

18. For the equilibrium:

$$2ICl(g) \rightleftharpoons I_2(g) + Cl_2(g)$$

the numerical value of the equilibrium constant K is $4{\cdot}8 \times 10^{-6}$ at 25°C.

Which of the following statements is true?

At 25°C at equilibrium

(A) there will be much less I_2 and Cl_2 than ICl.
(B) there will be twice as much ICl as I_2 present.
(C) the pressures of I_2, Cl_2, and ICl will all be the same.
(D) there will be much more I_2 and Cl_2 than ICl.

1997 HSC Q2

19. The effect of a change on an equilibrium in a chemical reaction can be predicted by Le Châtelier's principle. According to this principle, if a system at equilibrium is subject to change:

(A) There will be an increase in both the forward and reverse reactions to the same extent and so no change in the position of the equilibrium.
(B) The system will establish a new equilibrium position by increasing the rate of the reverse reactions.
(C) This will result in the system adjusting itself to return to equilibrium by partially counteracting the effect of the change
(D) The system will establish a new equilibrium position by increasing the rate of the forward reactions.

20. The numerical value of K_a for HCN is $6{\cdot}17 \times 10^{-10}$. The pH of a $0{\cdot}100$ mol L^{-1} solution of HCN is

(A) 1·00
(B) 5·10
(C) 6·17
(D) 9·21

1996 HSC Q12

21. Consider the reaction:

$$3Fe(s) + 4H_2O(g) \rightleftharpoons 4H_2(g) + Fe_3O_4(s)$$

When the total pressure is increased,

(A) more H_2 is produced.
(B) more H_2O is produced.
(C) no change occurs.
(D) more Fe is produced.

1997 HSC Q5

22. If the forward reaction of an equilibrium reaction is exothermic, what will happen to the position of an equilibrium system and its K_{eq} following an increase in the system's temperature?

(A) It will move to the right and K_{eq} will increase.
(B) It will move to the left and K_{eq} will increase.
(C) It will move to the right and K_{eq} will decrease.
(D) It will move to the left and K_{eq} will decrease.

23. Consider the reaction at equilibrium at 1000°C:

$$2CO(g) + O_2(g) \rightleftharpoons 2CO_2(g) \qquad \Delta H = -566 \text{ kJ mol}^{-1}$$

Which change would result in a larger concentration of CO_2?

(A) Decreasing the volume
(B) Increasing the temperature
(C) Adding a catalyst
(D) Decreasing the partial pressure of CO(*g*)

2000 HSC Q5

24. Which of the following is true for a system at equilibrium?

(A) The number of collisions per unit time between reactants is equal to the number of collisions per unit time between products.

(B) The product of the concentrations of the reactants is equal to the product of the concentrations of the products.

(C) Reactants are reacting to form products at the same rate as products are reacting to form reactants.

(D) All concentrations of reactants and products are equal.

1997 HSC Q13

25. Consider the reaction

$$FeO(s) + CO(g) \rightleftharpoons Fe(s) + CO_2(g) \qquad \Delta H = -283 \text{ kJ}$$

A change in conditions that moves the equilibrium position of this system to the right, i.e. favouring the products, is

(A) an increase in pressure.

(B) a decrease in temperature

(C) the addition of more finely powdered FeO.

(D) an increase in the concentration of carbon dioxide.

26. Addition of a catalyst may cause any except one of the following changes to occur. Identify the exception below.

(A) Change the value of the equilibrium constant for the reaction.

(B) Allow a new reaction mechanism having a lower activation energy.

(C) Increase the percentage of collisions that result in reaction.

(D) Increase the rates of both forward and reverse reactions.

27. Consider the equilibrium reaction by the equation below.

$$2NO_2(g) \rightleftharpoons N_2O_4(g) \qquad \Delta H = -60 \text{ kJ}$$

If the temperature is increased (volume kept constant), then the concentration of $N_2O_4(g)$ will

(A) remain the same because only the rates of the reaction change

(B) decrease because the formation of more $NO_2(g)$ absorbs energy

(C) increase because the reaction is exothermic

(D) decrease because the pressure will also rise.

28. An experiment is carried out to investigate the effect of temperature change on the reaction represented by the equation:

$$N_2O_4(g) \rightleftharpoons 2NO_2(g) \qquad \Delta H = 54{\cdot}8 \text{ kJ mol}^{-1}$$

What will result if the temperature increases?

(A) The value of the equilibrium constant will remain the same, but equilibrium will be reached more quickly.

(B) The value of the equilibrium constant will remain the same, but equilibrium will be reached more slowly.

(C) The value of the equilibrium constant will increase.

(D) The value of the equilibrium constant will decrease.

2000 HSC Q8

29. Which of the following pairs of solutions will form a precipitate?

(A) $Ba(NO_3)_2$ and Na_2SO_4

(B) KCl and $Ba(NO_3)_2$

(C) $NaCl$ and $Ba(NO_3)_2$

(D) Na_2SO_4 and KNO_3

30. What is the K_{sp} expression when lead(II) chloride dissolves in water to form a saturated solution?

(A) $[Pb^{2+}]^2\,[Cl^-]$

(B) $[Pb^{2+}]\,[Cl^-]^2$

(C) $[PbCl_2]$

(D) $\dfrac{[Pb^{2+}]\,[Cl^-]}{[PbCl_2]}$

31. What is the process of an ionic compound dissolving in water and separating into positive and negative ions called?

(A) Precipitation

(B) Equilibrium

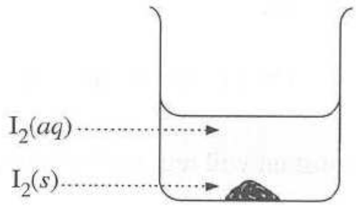
(C) Dissociation

(D) Solubility

32. Iodine is a solid that forms a brown solution in water.

$$I_2(s) \rightleftharpoons I_2(aq)$$

The system is at equilibrium.



Adding more solid iodine will

(A) make the solution darker brown.

(B) make the solution lighter brown.

(C) have no effect on the equilibrium.

(D) result in an increase in the concentration of $I_2(aq)$.

1999 HSC Q10

33. In ultraviolet light, fluorine molecules dissociate into fluorine atoms.

$$F_2(g) \rightleftharpoons 2F(g)$$

The concentrations at equilibrium are

$$[F_2] = 1{\cdot}0 \times 10^{-2} \text{ mol L}^{-1}$$
$$[F] = 3{\cdot}0 \times 10^{-4} \text{ mol L}^{-1}$$

What is the correct value for the equilibrium constant?

(A) $9{\cdot}0 \times 10^{-6}$

(B) $3{\cdot}0 \times 10^{-2}$

(C) $6{\cdot}0 \times 10^{-2}$

(D) $1{\cdot}1 \times 10^{5}$

1999 HSC Q5

34. According to collision theory, the chance of a reaction occurring depends on the following:

(A) The reactant particles colliding at an effective orientation.

(B) The frequency of the collisions between the reactant particles.

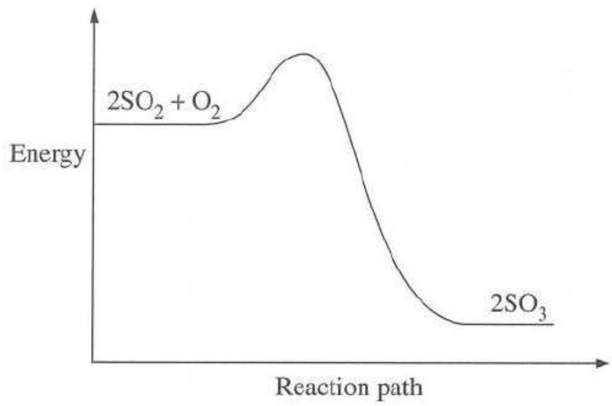
(C) Whether the energy between particles in a collision is equal to, or greater than the activation energy.

(D) All of the above.

35. Sulfur trioxide (SO_3) is used industrially to prepare sulfuric acid. It is formed by combining sulfur dioxide (SO_2) with an excess of air. Vanadium pentoxide (V_2O_5) may be used as a catalyst for this reaction.

$$2SO_2(g) + O_2(g) \xrightleftharpoons{V_2O_5} 2SO_3(g)$$

The energy profile for this reaction is shown below.



Which of the following would alter the equilibrium in favour of the formation of sulfur trioxide AND also increase the rate of reaction?

(A) Decreasing the volume of the reaction vessel at constant temperature.

(B) Increasing the temperature of the reaction vessel at constant pressure.

(C) Decreasing the temperature of the reaction vessel at constant pressure.

(D) Increasing the amount of vanadium pentoxide.

1999 HSC Q15

36. The brown gas nitrogen dioxide (NO_2) generally exists in equilibrium with the colourless gas dinitrogen tetroxide (N_2O_4).

$$\underset{\text{(brown)}}{2NO_2(g)} \rightleftharpoons \underset{\text{(colourless)}}{N_2O_4(g)} \qquad \Delta H^\circ = -54{\cdot}8 \text{ kJ mol}^{-1}$$

Which of the following changes would cause an increase in the equilibrium concentration of $N_2O_4(g)$?

(A) Adding nitrogen (N_2) gas.

(B) Introducing a catalyst.

(C) Decreasing the temperature in the container.

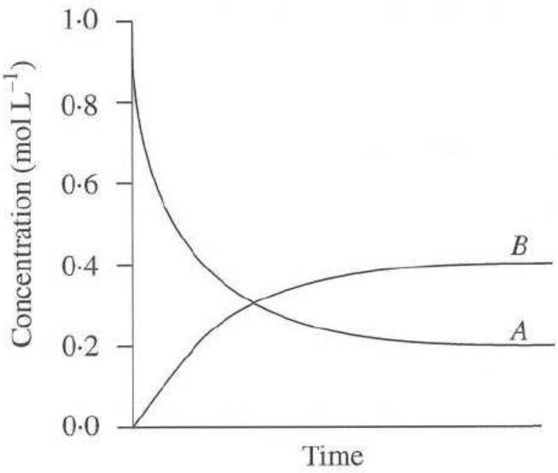
(D) Increasing the volume of the container.

1998 HSC Q11

37. If compound A is heated, it decomposes according to the equation:

$$2A(g) \rightleftharpoons B(g) + C(g)$$

The following diagram shows the progress of the reaction.



What is the equilibrium constant for the reaction?

(A) 0·8

(B) 2·0

(C) 4·0

(D) 10·0

2000 HSC Q15

Short-answer questions

Question 38 (4 marks)

Consider this chemical system which is at equilibrium.

$$X(g) + Y(g) \rightleftharpoons Z(g) + \text{heat}$$

(a) Explain the effect of decreasing the volume of the reaction vessel. **2**

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(b) Explain the effect of adding a catalyst to this equilibrium mixture. **2**

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2013 HSC Q24

Question 39 (3 marks)

Explain the impact of an increase in pressure and an increase in temperature on the solubility of carbon dioxide in water. **3**

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq) \qquad \Delta H \text{ is negative (exothermic)}$$

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Adapted 2012 HSC Q23

Question 40 (5 marks)

Hydrogen iodide is a colourless gas that will decompose into colourless hydrogen gas and purple iodine gas according to the following endothermic reaction.

$$2HI(g) \rightleftharpoons H_2(g) + I_2(g)$$

(a) A 1.0 L glass container was filled with 0.60 moles of hydrogen iodide gas. When equilibrium was established, there were 0.25 moles of iodine gas present in the container. **3**

Calculate the equilibrium constant for this reaction.

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(b) The container was then cooled. **2**

Explain the change in the appearance of its contents.

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2013 HSC Q32(b)(i) & (ii)

Question 41 (5 marks)

The equilibrium constant expression for a gaseous reaction is as follows:

$$K = \frac{[N_2][O_2]}{[NO]^2}$$

(a) Write the equation for this reaction. **1**

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Question 41 continues

Question 41 (continued)

(b) 0.400 moles of NO was placed in a 1.00 L vessel at 2000°C. The equilibrium concentration of N_2 was found to be 0.198 mol L^{-1}. **3**

Calculate the equilibrium constant for this reaction and use this value to describe the position of the equilibrium.

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(c) What could be changed that would result in a different value of K for this equilibrium? **1**

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End of Question 41

2012 HSC Q34(b)(i)–(iii)

Question 42 (4 marks)

The Haber process is shown by the equation: **4**

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H \text{ is negative}$$

For this reaction, a reaonably high temperature around 400–500°C is used and a relatively large pressure of about 250 atmospheres. Using Le Châtelier's principle, justify the choice of temperature and pressure conditions used to optimise the yield in this process.

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Adapted 2008 HSC Q23

Question 43 (6 marks)

Models are often used to help explain complex concepts.

You performed a first-hand investigation to model a dynamic equilibrium reaction.

(a) Outline the procedure used and the results you obtained. **2**

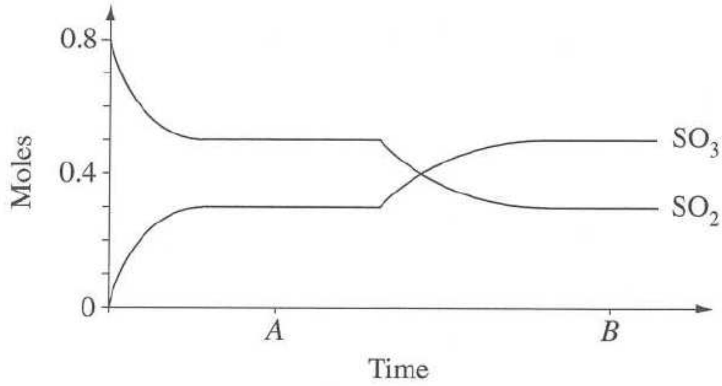
(b) Identify a risk associated with this procedure. **1**

(c) Describe how this procedure models equilibrium and state a limitation of the model. **3**

Adapted 2007 HSC Q28(e)(i)-(iii)

Question 44 (5 marks)

At room temperature 0.80 moles of SO_2 and 0.40 moles of O_2 were introduced into a sealed 10 L vessel and allowed to come to equilibrium.



(a) Write the equilibrium constant expression and calculate the value for the equilibrium constant at time *A*. **3**

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(b) Explain why a new equilibrium position was established at time *B*. **2**

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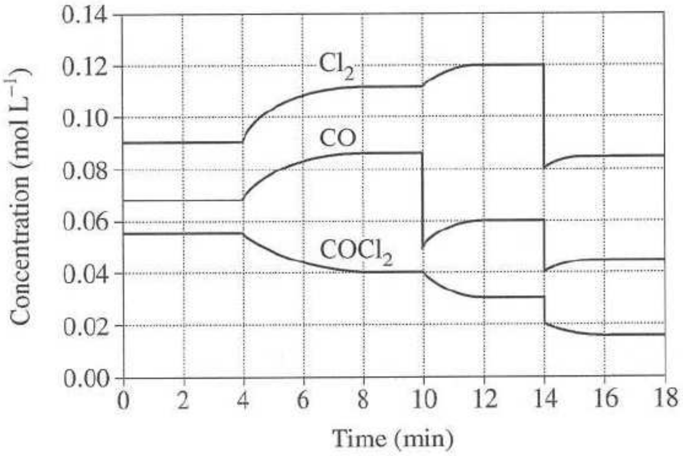
2010 HSC Q32(c)(i) & (ii)

Question 45 (6 marks)

The graph shows the variation in concentration of reactant and products as a function of time for the following system.

$$COCl_2(g) \rightleftharpoons Cl_2(g) + CO(g) \quad \Delta H = +108 \text{ kJ}$$



Identify and explain the changes in conditions that have shaped the curves during the time that the system was observed:

(a) initially: **2**

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(b) at 10 minutes: **2**

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(c) at 14 minutes: **2**

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Adapted 2009 HSC Q23

Question 46 (6 marks)

At a particular temperature, iodine trichloride dissociates into iodine gas and chlorine gas according to the following equation:

$$2ICl_3(g) \rightleftharpoons I_2(g) + 3Cl_2(g) \quad \Delta H = 240 \text{ kJ}$$

Initially 0.35 mol of $ICl_3(g)$ was introduced into a 1.0 L container and allowed to come to equilibrium. At equilibrium there was 0.45 mol L^{-1} of $Cl_2(g)$.

(a) Write the equilibrium constant expression for this reaction. **1**

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(b) Calculate the value of K at this temperature. **3**

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(c) What are TWO consequences of increasing the temperature of the mixture at equilibrium? **2**

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2009 HSC Q27(b)(i)-(iii)

Question 47 (8 marks)

Hydrogen sulfide can be removed from natural gas via the following process.

$$2H_2S(g) + SO_2(g) \rightleftharpoons 3S(s) + 2H_2O(g) \qquad \Delta H = -145 \text{ kJ mol}^{-1}$$

(a) Write the equilibrium constant expression for this reaction. **1**

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Question 47 continues

Question 47 (continued)

(b) Calculate the equilibrium constant, when 1.00 mol of H_2S and 1.00 mol of SO_2 react in a 1.00 L vessel at 373 K to give 0.50 mol of water vapour under equilibrium conditions. **3**

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(c) Identify FOUR factors that would maximise the removal of $H_2S(g)$ in this reaction. **4**

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End of Question 47

Adapted 2007 HSC Q28(c)(i)-(iii)

Question 48 (4 marks)

Methane can be produced by the following reaction:

$$CS_2(g) + 4H_2(g) \rightleftharpoons CH_4(g) + 2H_2S(g)$$

(a) Write the expression for the equilibrium constant for this reaction. **1**

....

(b) At equilibrium, a 5·0 L vessel contains 0·55 mol CH_4, 0·125 mol H_2S, 0·15 mol CS_2, and 0·15 mol H_2. **3**

Calculate the value of the equilibrium constant for this reaction at the temperature of the experiment.

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1996 HSC Q20

Question 49 (6 marks)

Consider the following mixture of gases in a closed 5.0 L vessel at 730°C.

Gas	*Quantity* (mol)	*Concentration* (mol L^{-1})
CH_4	2.00	
H_2O	1.25	
CO	0.75	
H_2	0.75	

The following reaction occurs:

$$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g) \qquad \Delta H = +206 \text{ kJ}$$

The equilibrium constant, K, is 0.26 at 730°C.

(a) Complete the table above, then determine whether the system is at equilibrium. **3**

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(b) Explain how conditions in this reaction could be adjusted to increase the quantity of products. **3**

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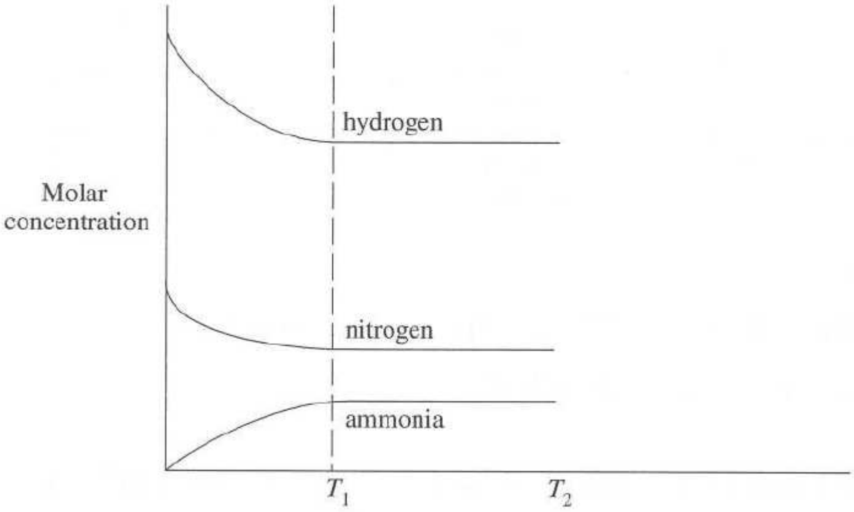
2008 HSC Q29(c)(i) & (ii)

Question 50 (5 marks)

The graph shows the variation in concentrations of reactants and product with time for the Haber process. This process is shown by the equation:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H \text{ is negative}$$



(a) State why the concentrations of reactants and product do not change between T_1 and T_2. **1**

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(b) At time T_2 the volume of the reaction vessel was reduced.

(i) Sketch on the graph how the concentrations of reactants and product would change after the volume was reduced. **2**

(ii) Explain the changes shown on your graph. **2**

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Adapted 2005 HSC Q27

Question 51 (6 marks)

Nitrogen dioxide forms an equilibrium mixture with dinitrogen tetraoxide as shown.

$$2NO_2(g) \rightleftharpoons N_2O_4(g)$$

At 100°C, K for this reaction is 2.08.

At 25°C, a 1.00 L vessel initially contained 0.132 mol of $NO_2(g)$. Once equilibrium had been established, there was 0.0400 mol of $N_2O_4(g)$ in the vessel.

(a) Explain the effects of the addition of a catalyst and an increase in pressure on the yield of N_2O_4 in this reaction when carried out at 25°C. **2**

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(b) Calculate the equilibrium constant for this reaction at 25°C, and account for any difference from the K value at 100°C. **4**

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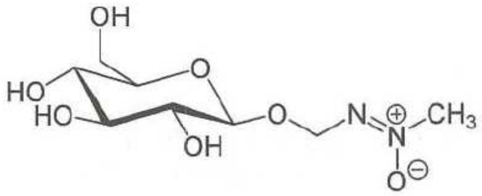
Adapted 2005 HSC Q28(b)(i) & (ii)

Question 52 (9 marks)

Cycad seeds contain the harmful cycasin toxin. This cycasin toxin has a solubility in water of 56.6 g L^{-1}. Its structure is shown below.



In the past, Aboriginal and Torres Strait Islander Peoples developed an understanding about how to prepare cycad seeds, so they could be safely used as a source of food. One method used was to firstly remove the non-edible, fleshy outer layer (sarcotesta) – this is often wrongly confused as being a fruit. Then, although they had no understanding of the science behind what they were doing, they worked out how to detoxify the seed and make it edible. This was done by placing the cycad seeds in the running water of a creek or in a lake, so that the toxin was removed.

(a) The cycad seeds were initially either cooked or pounded before leaching them in water. What would have been achieved by doing this? **2**

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(b) The scientific method being used to remove the toxin is known as *leaching*. Outline the process of leaching and the property of cycasin that allowed it to be leached from the cycad seeds. **2**

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Question 52 continues

Question 52 (continued)

(c) Table sugar, sucrose, is very water soluble, as about 2 kg will dissolve in a litre of water at 20°C, whereas cycasin has a solubility in water of 56.6 g L^{-1}. **3**

Explain, in terms of solubility and equilibria, why would it be preferable to place cycad seeds in running water for a period of time, rather than soaking them in a wooden bowl.

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(d) Outline whether the leaching method used by Aboriginal and Torres Strait Islander Peoples to detoxify cycad seeds indicates that they had an understanding of the science of solubility equilibria or not. **2**

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End of Question 52

Question 53 (8 marks)

Consider the equilibrium system:

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons [FeSCN]^{2+}(aq)$$

In this reaction, iron(III) ions react with the colourless thiocyanate (SCN^{-}) to form a deep blood-red coloured solution due to the formation of iron(III) thiocyanate ions.

(a) Name a quantitative technique that could be used to measure the colour of the reactants and products in order to determine K_{eq} for this reaction. **1**

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(b) Write the equilibrium constant expression for this reaction. **1**

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(c) Indicate whether the blood-red colour of the solution would decrease / increase if the following are added:

(i) more $SCN^{-}(aq)$ are added to the solution? **1**

(ii) a small amount of $Fe(NO_3)_3$ is added? **1**

(iii) a small amount of $AgNO_3$ solution (so that insoluble AgSCN forms)? **1**

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(d) Using Le Châtelier's principle, explain what happen if a small amount of water was added to the solution. **3**

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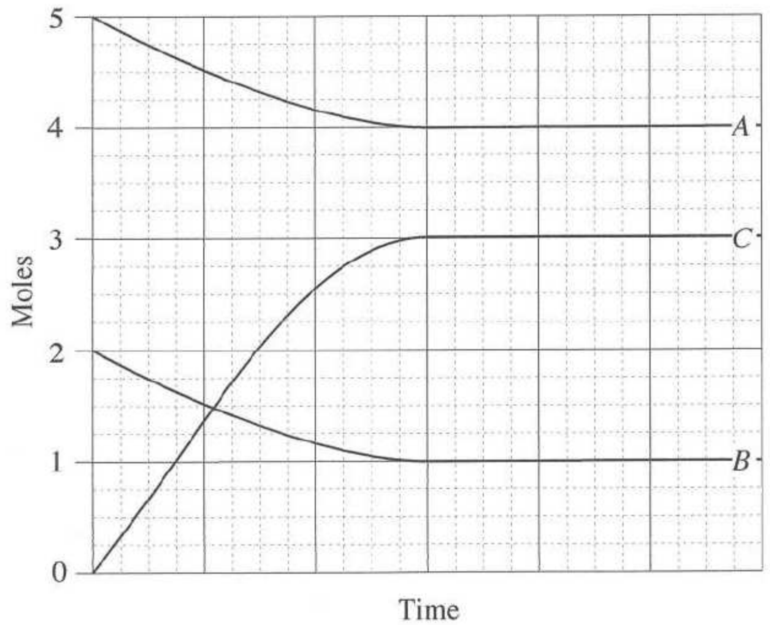
Question 54 (3 marks)

Two substances A and B react according to the equation:

$$A + B \rightleftharpoons 3C$$

5 moles of A and 2 moles of B are mixed in a 2 L closed container. The reaction is allowed to come to equilibrium at temperature T. The graph below shows the variation in moles of A, B, and C over time.



Calculate the value of the equilibrium constant, K, for the reaction at temperature T. Show your working. **3**

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1996 HSC Q17

Question 55 (3 marks)

The colourless gas nitrogen monoxide (NO) reacts with oxygen to form a brown gas nitrogen dioxide (NO_2). This results in a light-brown equilibrium mixture according to the equation:

$$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g) \qquad \Delta H = -5{\cdot}66 \text{ kJ}$$

What change in colour of this equilibrium mixture is observed if:

(a) the pressure is decreased?

.... 1

(b) more oxygen gas is added?

.... 1

(c) a catalyst is added?

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1995 HSC Q22

Question 56 (3 marks)

When solid AgBr is added to water, an equilibrium between the solid and dissolved ions is established, as follows: 3

$$AgBr(s) \rightleftharpoons Ag^{+}(aq) + Br^{-}(aq)$$

The solubility product (K_{sp}) for AgBr is 5.0×10^{-13}.

What is the concentration of Ag^+ in a saturated solution of AgBr?

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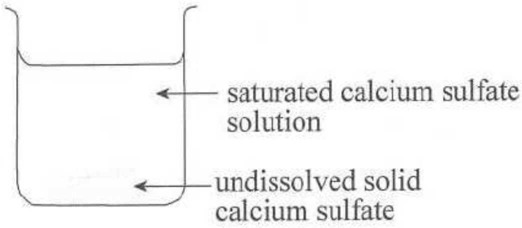
Question 57 (8 marks)

The diagram represents a beaker containing a saturated solution of calcium sulfate. There is a small amount of undissolved solid calcium sulfate at the bottom of the beaker.



(a) Write the equation that represents the equilibrium in the beaker involving solid calcium sulfate and its ions. **1**

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(b) Write the expression for the solubility product K_{sp} of calcium sulfate **1**

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(c) The solubility of calcium sulfate in water is 0.0045 mol L^{-1} at 25°C. Calculate the value of K_{sp} of calcium sulfate at 25°C. Show working in your answer. **2**

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(d) Would the solubility of calcium sulfate in a 0.010 mol L^{-1} solution of sodium sulfate be different from its solubility in water? Give a reason for your answer. **4**

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Adapted 1993 HSC
Q El 3(a)(i)-(iii), (v)

Question 58 (8 marks)

Silver chromate (Ag_2CrO_4), a yellow solid, dissolves in water to give a yellow saturated solution containing 2.5×10^{-2} g L^{-1} silver chromate. It is also known that chromate ions react with hydrogen ions as follows:

$$\underset{\textit{yellow}}{2CrO_4^{2-}(aq)} + 2H^+(aq) \rightleftharpoons 2HCrO_4^-(aq) \rightleftharpoons \underset{\textit{orange}}{Cr_2O_7^{2-}(aq)} + H_2O(l)$$

(a) Calculate the molar concentration of chromate ions in a saturated solution of silver chromate in pure water. Your answer needs to include an equation that represents silver chromate dissolving in water. **2**

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(b) Calculate K_{sp} for silver chromate. **2**

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(c) Calculate the molar concentration of chromate ions in a saturated solution of silver chromate in 0.10 mol L^{-1} silver nitrate. **2**

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(d) Explain why silver chromate more readily dissolves in dilute nitric acid than in pure water. **2**

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Adapted 1987 HSC Q El 3(d)(i)–(iv)

Question 59 (5 marks)

The solubility products (K_{sp}) for two silver salts are given below.

silver chloride, $AgCl$ $K_{sp} = 2.0 \times 10^{-10}$

silver carbonate, Ag_2CO_3 $K_{sp} = 8.0 \times 10^{-12}$

(a) Which silver salt has the greater solubility (mol L^{-1})? **3**

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(b) Calculate the solubility (mol L^{-1}) of $AgCl$ in a 0.10 mol L^{-1} solution of sodium chloride. **2**

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Adapted 1992 HSC Q El 3(c)(i) & (ii)

Question 60 (2 marks)

Photosynthesis by algae and other plant life in sunny weather can lead to higher than normal pH values in rivers. This is due to a disturbance in the carbonate/bicarbonate equilibrium. Use the equation below for this equilibrium to explain how a rise in pH could occur. **2**

$$OH^- + CO_2 \underset{}{\overset{H_2O}{\rightleftharpoons}} HCO_3^- \overset{H_2O}{\rightleftharpoons} H^+ + CO_3^{2-}$$

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1997 HSC 35(c)

Question 61 (5 marks)

Initially, a mixture has 0·400 mol each of CO_2 and H_2 in a 1·00 L vessel kept at 980°C.

The reaction is $CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g)$

and at this temperature the equilibrium concentration of CO is 0·225 mol L^{-1}.

(a) What is the equilibrium concentration of H_2? **2**

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(b) Write the equilibrium constant expression for this reaction. **1**

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(c) Calculate the equilibrium constant at 980°C. **2**

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Adapted 1998 HSC Q23

Question 62 (5 marks)

Phosphorus trichloride (PCl_3) reacts with chlorine (Cl_2) to produce phosphorus pentachloride (PCl_5).

$$PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g)$$

Consider this system at equilibrium.

(a) Compare the rate of the forward reaction with the rate of the reverse reaction. **1**

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Question 62 continues

Question 62 (continued)

(b) If the gases are in a sealed container and the volume is suddenly increased, explain any shift in the *equilibrium position*. **2**

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(c) If the temperature is kept constant as more $Cl_2(g)$ is added, what effect will this have on the *equilibrium constant*? **2**

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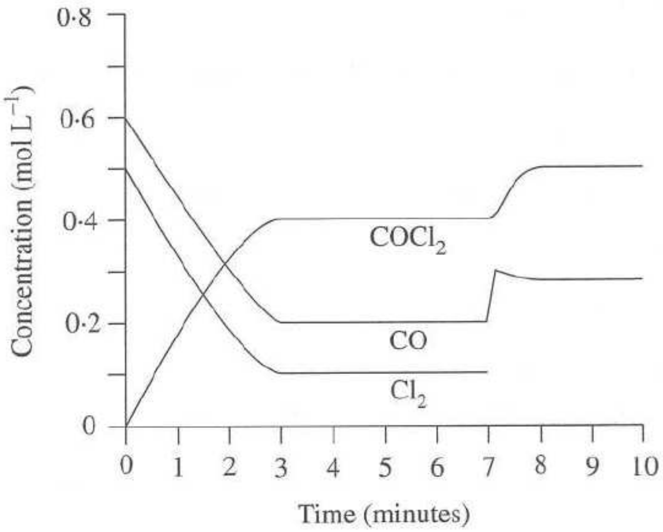
End of Question 62 *1999 HSC Q18*

Question 63 (4 marks)

Phosgene is produced from chlorine and carbon monoxide according to the equation:

$$CO(g) + Cl_2(g) \rightleftharpoons COCl_2(g)$$

When CO and Cl_2 are mixed in the presence of activated carbon, the concentrations of each gas change according to the graph below.



Question 63 continues

Question 63 (continued)

(a) Calculate a value for the equilibrium constant. **2**

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(b) The change in concentration of chlorine after 7 minutes is not shown on this graph. **2**

(i) What was added at 7 minutes? ....

(ii) Predict the effect of this addition on the concentration of Cl_2. A calculation is not required.

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End of Question 63

1999 HSC Q24

Question 64 (3 marks)

The production of hydrogen iodide from hydrogen and iodine is given by the equation:

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

1·0 mol hydrogen and 1·0 mol iodine were introduced into a sealed 1 L reaction vessel at 500°C and allowed to come to equilibrium. It was then found that 1·55 mol hydrogen iodide had been produced.

(a) Calculate the equilibrium concentrations of H_2 and I_2 under these conditions. **2**

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(b) Calculate the value of the equilibrium constant (K) for the reaction. **1**

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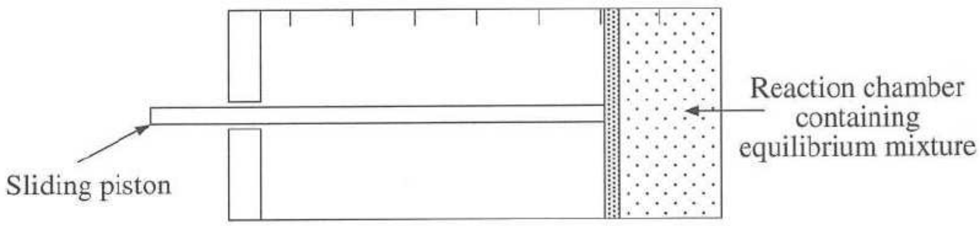
2000 HSC Q25

Question 65 (3 marks)

A student investigated the equilibrium system involving brown nitrogen dioxide (NO_2) and colourless dinitrogen tetroxide (N_2O_4).

$$\underset{\text{brown}}{2NO_2(g)} \rightleftharpoons \underset{\text{colourless}}{N_2O_4(g)}$$

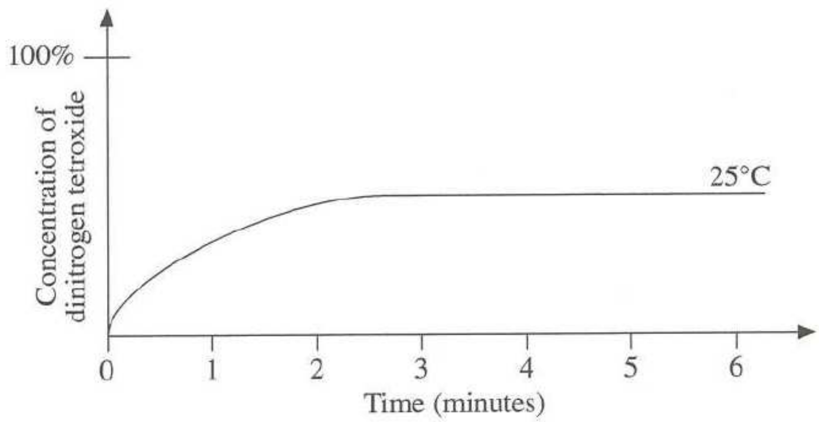
Nitrogen dioxide gas was placed into a reaction chamber as shown and allowed to reach equilibrium.



The colour of the equilibrium mixture was compared under different reaction conditions.

Temperature (°C)	*Pressure* (kPa)	*Colour*
0	1·0	light brown
25	1·0	brown
100	1·0	dark brown

The graph represents the change in concentration of N_2O_4 as the reaction proceeds to equilibrium at 25°C.

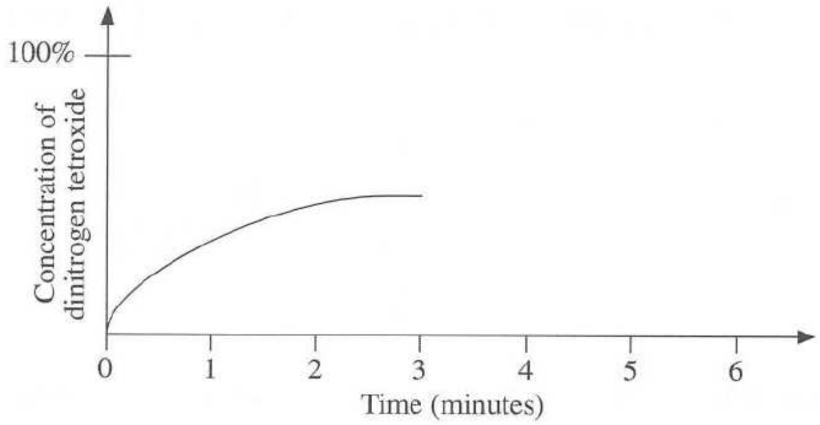


Question 65 continues

Question 65 (continued)

(a) On the graph above, sketch another curve to show the change in concentration of N_2O_4 as the reaction proceeds to equilibrium at 100°C. **1**

(b) On the graph below, continue the curve to show the change in concentration of N_2O_4 if the piston had been compressed after three minutes. **2**



End of Question 65

Adapted 1999 HSC Q21

Question 66 (5 marks)

The table shows the effect of temperature on the equilibrium constant (K) for the reaction:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

Temperature (°C)	K
700	2·63
800	0·915
900	0·384
1000	0·184
1100	0·098

Question 66 continues

Question 66 (continued)

(a) Plot the data on the grid below. Include a curve of best fit to show the trend clearly. **3**

(b) Is the reaction endothermic or exothermic? Give a reason for your answer. **2**

End of Question 66 *2000 HSC Q19*

Question 67 (4 marks)

The solution of limestone, $CaCO_3$, depends on the action of carbon dioxide and water.

$$CaCO_3(s) + H_2O(l) + CO_2(g) \rightleftharpoons Ca^{2+}(aq) + 2HCO_3^-(aq) \qquad \Delta H = -42\ kJ$$

Deep in limestone caves, the concentration of carbon dioxide is often greater than the concentration of carbon dioxide in the atmosphere.

(a) What effect would the reduction of concentration alone have on the equilibrium as a river leaves a limestone cave? Explain your answer. **2**

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(b) Name ONE other change which could occur to alter the equilibrium concentration of calcium ions as a river leaves a limestone cave. Explain your answer. **2**

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Adapted 1984 HSC Q14

Answers

Module 5: Equilibrium and Acid Reactions

Multiple choice

1. C	**2.** C	**3.** C	**4.** B	**5.** C	**6.** D	**7.** D	**8.** A
9. A	**10.** A	**11.** C	**12.** A	**13.** D	**14.** D	**15.** D	**16.** A
17. C	**18.** A	**19.** C	**20.** B	**21.** C	**22.** D	**23.** A	**24.** C
25. B	**26.** A	**27.** B	**28.** C	**29.** A	**30.** B	**31.** C	**32.** C
33. A	**34.** D	**35.** A	**36.** C	**37.** C			

Explanations

1. **C** ΔH is positive so the reaction to produce $NO_2(g)$ is endothermic. Therefore, raising the temperature will result in increased $[NO_2(g)]$ as the equilibrium shifts towards the right to use up heat energy as predicted by Le Châtelier's Principle. So (C) is the answer. A catalyst increases the rate of reaction, not the yield – so (A) is incorrect. Decreasing the volume will reduce the yield, not increase it – so (B) is incorrect. Adding argon will not alter the yield, so (D) is incorrect.

2. **C** If the reaction is endothermic, cooling the solution will favour the reverse reaction, so it will turn pink, as in (C). Diluting the solution with water will favour the reverse reaction, so it will turn pink, not blue. So (A) is incorrect. If the reaction is exothermic, heating the solution will favour the reverse reaction, so it will turn pink, not blue. So (B) is incorrect. Adding KCl solid to the solution will favour the forward reaction, so it will turn blue, not pink. So (D) is incorrect.

3. **C** The consequences of carrying out the reaction at 400°C compared to 200°C are:

- reaction rate at 400°C will be faster than at 200°C. Steeper curves in (A), (C) and (D) show this. (B) is incorrect as it has a slower rate of NH_3 production.
- yield of ammonia at 400°C will be less than at 200°C, because the forward reaction (formation of NH_3) is exothermic. Hence, lower temperatures favour the formation of NH_3 (Le Châtelier's Principle). Only the graph in (C) shows the lower yield, and so is the answer. So (A) and (D) are incorrect.

4. **B** An increase in products with increased pressure will occur if the reaction moves to the right. The only reaction where there are more gaseous molecules on the left for this to occur is in (B). So (B) is the answer.

5. **C** If the temperature is lowered, an equilibrium will move in the direction that tends to increase temperature (i.e. release heat). As this reaction is exothermic, the reaction will move to the right if the temperature is lowered. If the pressure is increased, an equilibrium will move in the direction that tends to reduce pressure (i.e. to a decrease in the number of moles of gas). High pressure will favour the forward reaction in this case as two moles of gas are reacting and one mole of gas is being formed. Only (C) has this correct combination.

6. **D** Adding more reactant will shift the equilibrium to the right, so (D) is the answer. In equilibrium reactions, the addition of a catalyst will speed up the reaction, but does not shift its equilibrium position to the left or right, so (A) is incorrect. Pressure only effects equilibrium if there are different numbers of moles of gases in the reactants compared to the products. There are 2 moles of gas on each side in this case, so (B) is incorrect. The reaction is exothermic, so increasing the temperature will shift the equilibrium to the left not the right, so (C) is incorrect.

7. **D** Le Châtelier's principle can be used to predict how the equilibrium will adjust if the concentration of chemical substances that are in equilibrium are changed in any way, so (D) is the answer. Le Châtelier's principle does not predict products, nor the rate at which this equilibrium is attained, so (A) and (B) are incorrect. Catalysts do not affect the equilibrium position, so (C) is incorrect.

8. **A** An equilibrium system needs to be a closed system, which can only exchange energy with its surroundings, as in (A). It must not be an open system which can exchange energy, as well as matter with its surroundings. Also, an equilibrium can only occur with a reversible reaction, in which both the forward and reverse reactions can occur. So (B), (C) and (D) are incorrect.

9. **A** For the time prior to T_1, the graph shows the reactants being depleted as a solid line and a line with 'small' dashes. It shows the product being formed as a line with 'long' dashes. The equation shows that more $H_2(g)$ is used in the reaction than $CO(g)$, so $H_2(g)$ is the reactant that is depleted more quickly (= line with 'small' dashes). The graph shows that at time T_1, additional $H_2(g)$ is added. At time T_2, some of the product CH_3OH is removed. (A) is the only possible answer.

10. A From Le Châtelier's Principle, a decrease in pressure will cause the equilibrium to shift to counter the decrease in pressure. So the equilibrium will be disturbed, so (B) is incorrect. To counter the decrease in pressure, the equilibrium will shift to the side with more gaseous molecules, so the reverse reaction rate will increase. Hence the concentration of reactants will increase, and the concentration of products will decrease. So (C) and (D) are incorrect. The reaction is exothermic, so the reverse reaction will be endothermic (absorb heat), so (A) is the answer.

11. C H^+ is a reactant. A change in the pH of the reaction mixture will change $[H^+]$ and therefore affect the position of the equilibrium. Sodium acetate is the salt of a strong base and a weak acid, so the addition of this salt will result in acetate ions removing H^+ to form acetic acid. This causes the equilibrium to shift to the left producing more $CrO_4^{2-}(aq)$. So (C) is the answer. Sodium nitrate and sodium chloride are both salts of a strong acid and a strong base, so their addition will not affect pH and so would not affect the equilibrium. So (A) and (B) are incorrect. The addition of ammonium chloride, the salt of a weak base and a strong acid, will decrease pH as $[H^+]$ increases. This will shift the equilibrium to the right, thus reducing $[CrO_4^{2-}(aq)]$. So (D) is incorrect.

12. A ΔH is negative, i.e. the reaction is exothermic. Since all species are gases, Le Châtelier's principle indicates that if you increase the pressure, the system will respond by favouring the reaction that produces fewer molecules, i.e. higher pressure favours the production of ammonia. So (C) and (D) are incorrect as they show a decrease in ammonia. Since this reaction is exothermic, it will produce a higher yield at lower temperatures, according to Le Châtelier's Principle. (B) is incorrect as it shows higher yields at higher temperatures, and (A) is the answer as it correctly shows higher yields at lower temperatures.

13. D Equilibrium is reached when the concentrations become constant as the rates of the forward and reverse reactions become equal and not when the reactant and product concentrations become equal as in (A). The forward reaction continues so (B) is wrong. Equilibrium is reached after 4 hours not 2 hours, so (C) is also wrong. (D) is correct as both reactants and products are present at equilibrium.

14. D The expression for an equilibrium constant, has the products over the reactants, therefore (A) and (C) are incorrect. (B) is incorrect as it does not correctly include the numbers of moles involved, whereas (D) correctly includes these.

[Note: this is easy to remember by the mnemonic PORK: Products Over Reactants = K]

15. D To reach equilibrium, all products and reactants must be present in a closed system as in (D). So, (A) and (C) are incorrect as they are in an open container. (B) is incorrect as there can be no equilibrium with only one substance present.

16. A A non-reversible reaction is where the reaction can only go in one direction, and so the products cannot be converted back to the reactants. An open system can exchange energy and matter with its surroundings, while a closed system does not allow any reactants or products to enter or escape. A closed system is necessary for a reversible reaction to reach equilibrium. Hence (A) is an incorrect statement and so is the answer.

17. C Think of Pork!!! **P**roducts **O**ver **R**eactants = ***K***.

∴ The bigger the value of K, the greater the tendency for the reaction to reach completion (the conditions are irrelevant to this question).

$5.0 \times 10^{-31} \ll 1.0 \times 10^{-13} \ll 4.7 \times 10^{-4} \ll 1.0 \times 10^{228}$

∴ the order must be IV, II, III, I ... as only in (C).

18. A To determine the answer, it is essential to have the correct equilibrium expression. Think of Pork!!! **P**roducts **O**ver **R**eactants = ***K***.

$K = \frac{[I_2(g)][Cl_2(g)]}{[ICl(g)]^2} = 4.8 \times 10^{-6}$ The smaller the value of K, the smaller the proportion of products compared to reactants. Since $K \lll 1.0$, $[ICl(g)]$ must be very much greater than either $[I_2(g)]$ or $[Cl_2(g)]$ as in (A), and not just twice as great as in (B). (C) and (D) are incorrect, because for (C) to be correct, $K = 1$, and for (D) to be correct, $K \ggg 1$.

19. C According to Le Châtelier's principle, if a system at equilibrium is subject to a change in conditions, then it will adjust itself to partially reduce the effect of the change. So (C) is the answer. (A) is incorrect as this is what happens when a catalyst is added. (B) and (D) are incorrect as they are what can happen when the temperature is changed, depending on whether the reaction is exothermic or endothermic.

20. B

$$K = \frac{[CN^-][H^+]}{[HCN]} = 6.17 \times 10^{-10}$$

Since $[H^+] = [CN^-]$ and since $[H^+] \lll [HCN]$

$$6.17 \times 10^{-10} = \frac{[H^+]^2}{[HCN]} = \frac{[H^+]^2}{0.100}$$

$$[H^+] = \sqrt{6.17 \times 10^{-10} \times 0.100} = 7.85 \times 10^{-6}$$

$$pH = 5.10$$ ... as in answer (B)

21. C Pressure changes will only affect the equilibrium if the reaction has a change in the number of moles of gaseous compounds between the reactants and the products. In this reaction, the equation shows $4H_2O(g)$ in the reactants and $4H_2(g)$ in the products so pressure will have no effect on the equilibrium as in (C). (A), (B) and (D) each involve a change in equilibrium so are all incorrect.

22. D Increasing the temperature for an exothermic reaction shifts the reaction to the left (i.e. the reverse direction) and decreases the K_{eq}. So (D) is the answer, and (B) is incorrect. Heating an endothermic reaction will shift the reaction to the right and increase the K_{eq}. So both (A) and (C) are incorrect.

23. A Decreasing the volume is equivalent to increasing the pressure. So, according to Le Châtelier's Principal, equilibrium will shift to the right increasing $[CO_2(g)]$. So (A) is the answer. As the forward reaction is exothermic, increasing temperature shifts the equilibrium to the left, as will decreasing the $[CO(g)]$. So, (B) and (D) are incorrect. Adding a catalyst, will not change the equilibrium concentrations at all, but will speed up the attainment of equilibrium, so (C) is incorrect.

24. C Equilibrium is reached when the concentrations become constant as the rates of the forward and reverse reactions become equal, as in (C) … and not when the reactant and product concentrations become equal, as in (D). (A) is clearly false as it is not the number of collisions that matters but the proportion of these collisions that result in chemical change. (B) is incorrect because it implies that all reactions would have the same equilibrium constant, equal to 1.

25. B Since this reaction is exothermic, a decrease in temperature will shift equilibrium to the right, i.e. to an increase in the products. So (B) is the answer. An increase in pressure will produce no change as there are the same number of gaseous molecules on both sides. So (A) is incorrect. The addition of more reactant in the form of powdered solid FeO will not shift it to the right, while the addition of more product, CO_2, will shift it to the left. So (C) and (D) are incorrect.

26. A A catalyst does not change the value of K_{eq} or the position of an equilibrium system. So (A) is the answer.

27. B Since this reaction is exothermic, an increase in temperature will shift it to the left, i.e. to more $NO_2(g)$ and therefore less product, i.e. less $N_2O_4(g)$. So only (B) is correct and therefore the answer.

28. **C** The forward reaction is endothermic, since ΔH is positive. So, increasing temperature will cause equilibrium to shift towards the right (Le Châtelier's Principal), increasing products and reducing reactants. So, the equilibrium constant (K) increases, as in (C) – and the higher temperature means that the new equilibrium will be reached more quickly. K will not be the same as in (A) and (B), or decreasing as in (D).

29. **A** Remember the solubility rules! Most sulfates are soluble (except Sr^{2+}, Ba^{2+}, Pb^{2+}). So $BaSO_4$ is insoluble. So (A) is the answer. All nitrates are soluble, all Na^+ and K^+ salts are soluble and most chlorides are soluble (except Ag^+, Pb^{2+}). So $BaCl_2$ is soluble. Hence (B) and (C) are not the answer. K_2SO_4 is soluble, so (D) is not the answer.

30. **B** Write a balanced chemical equation for $PbCl_2$ and its ions:

$$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)$$

Only aqueous species are in the K_{sp} and each species is raised to the power of its coefficient in the balanced equation. So $K_{sp} = [Pb^{2+}]\,[Cl^-]^2$ … as in (B).

[Note: Remember the K_{sp} expression only applies to 'insoluble' or 'sparingly soluble' salts, such as $PbCl_2$ as in this question!]

31. **C** Ionic substances dissolve by dissociation. So (C) is the answer.

32. **C** The system is at equilibrium, so the $I_2(aq)$ solution is saturated, i.e. no more iodine can dissolve. So, $[I_2(aq)]$ has already reached the maximum possible and its concentration will not change. So (D) is incorrect. There will be no colour change, so (A) and (B) are incorrect. Adding more $I_2(s)$ therefore has no effect on the equilibrium, as in (C).

33. **A** Remember PORK … **P**roducts **O**ver **R**eactants = ***K***

$$K = \frac{[F(g)]^2}{[F_2(g)]} = \frac{(3.0 \times 10^{-4})^2}{1.0 \times 10^{-2})} = 9.0 \times 10^{-6} \quad \text{… as in (A).}$$

34. **D** The basis of the collision theory is that a chemical reaction will occur when reactant particles collide, but not all collisions will produce a reaction. The chance of a chemical reaction occurring depends on all of the conditions given. So (D) is the answer.

35. **A** There are more gaseous molecules on the reactant side than on the product side, so decreasing the volume thereby increasing the pressure favours SO_3 production and because of an increase in the number of collisions also increases the reaction rate, so (A) is correct. The energy profile shows that the reaction is exothermic. Increasing the temperature as in (B) shifts the equilibrium towards reactants. So (B) is incorrect. A decrease in temperature as in (C) would favour SO_3 production, but decreasing the temperature reduces the reaction rate. So (C) is incorrect. The addition of catalyst increases the reaction rate, but does not change the equilibrium position. So (D) is incorrect.

36. **C** Adding $N_2(g)$ which is not involved in the reaction will not affect the equilibrium position, nor will the addition of a catalyst affect equilibrium position, so (A) and (B) are incorrect. The reaction left to right is exothermic, so decreasing the temperature will favour the formation of $N_2O_4(g)$. So (C) is the answer. Increasing the volume of the container as in (D) will immediately decrease the $N_2O_4(g)$ and the reaction will move towards the formation of $NO_2(g)$, leading to a further reduction in $N_2O_4(g)$. So (D) is incorrect.

37. **C**

	$2A(g)$	$\rightleftharpoons$	$B(g)$	$+$	$C(g)$
Mole ratio	2		1		1
$c_{initial}$	1 −		0 +		0 +
c_{end}	0.2		0.4		0.4
c_{change}	0.8		0.4		0.4

We can calculate the change in concentration (c_{change}) for A by subtracting 0.2 from 1 to give 0.8. As the mole ratio is 2 : 1 : 1, when 0.8 moles of A react, they form 0.4 moles of B and 0.4 moles of C.

$\therefore\ K = \frac{[products]}{[reactants]} = \frac{[B]\,[C]}{[A]^2} = \frac{(0.4 \times 0.4)^2}{(0.2)^2} = 4.0$ as in (C).

Short-answer questions

38. (a) Le Châtelier's Principle indicates that decreasing the volume will favour the forward reaction. So there will be an increase in the proportion of *Z*, as there are less moles of gaseous products than the number of moles of gaseous reactants.

(b) Since the reaction is at equilibrium, the catalyst will speed up both the forward and reverse reactions equally. So there will be no change in the proportions of *X*, *Y* and *Z* in the equilibrium mixture.

39. Each C=O bond in CO_2 is polar and so it reacts with water to form a weak acid. $CO_2(g)$ is the only gas in the reaction, so an increase in pressure will cause the equilibrium to shift to the right forming more H_2CO_3 to reduce the gas pressure, i.e. $CO_2(g)$ becomes more soluble in water.

Since the reaction is exothermic, if you increase the temperature, the equilibrium will shift in the endothermic direction. So, it moves to the left to form more $CO_2(g)$, i.e. $CO_2(g)$ becomes less soluble in water.

40. (a)

	$2HI(g)$	$\leftrightharpoons$	$H_2(g)$	+	$I_2(g)$
Initial concentration:	0.60		0		0 mol L^{-1}
At equilibrium:	$0.60 - (2 \times 0.25)$		0.25		0.25

$\therefore$ equilibrium constant, $K = \frac{[H_2][I_2]}{[HI]^2} = \frac{0.25 \times 0.25}{0.1^2} = 6.25$

(b) As the forward reaction is endothermic, the cooling process shifts the equilibrium for the reaction towards the formation of hydrogen iodide. This reduces the iodine concentration, so the colour intensity will decrease and become a paler purple.

41. (a) $2NO(g) \rightleftharpoons N_2(g) + O_2(g)$

(b)

0.4	0	0 mol L^{-1}	(initial conc)
$0.4 - (2 \times 0.198)$	0.198	0.198 mol L^{-1}	(equilibrium conc)

$K = \frac{[N_2][O_2]}{[NO]^2} = \frac{0.198 \times 0.198}{0.004^2} = 2450.25 \approx 2.45 \times 10^3$

The *K* value is a large number, so the equilibrium lies well on the products side – – indicating that the decomposition of nitric oxide is almost complete at 2000°C.

(iii) Temperature

42. The forward reaction is exothermic. Le Châtelier's principle indicates that when an increase in temperature occurs, the reverse reaction is favoured. So, the formation of ammonia would thus be favoured by lower temperatures. So the temperature would have been chosen so that it was not too high, but low enough to optimise the yield.

Le Châtelier's principle indicates that if the pressure is changed, the system will shift its equilibrium position to counteract this change. So, an increase in pressure will favour a shift toward the right with more products being formed. Hence the use of a relatively large pressure, as this would favour the formation of ammonia.

43. (a) PROCEDURE – People dancing in a room were used to model equilibrium. We started with the door shut, so no-one could leave/enter. The people had to join hands to dance, and split up to become single again. Then the doors were opened briefly to allow more people in and some to go out. We also closed off part of the room, so more people were pushed together to make couples.

RESULTS – When people ('reactants') joined up to dance in the closed room they became 'products'. When they split up and stopped dancing, they became 'reactants' again. When the doors were opened briefly and people went in and out, the number of 'reactants' and 'products' changed. When part of the room was closed off, the number of 'reactants' and 'products' changed again.

[Note: Equilibrium can also be modelled using coloured beads that connect/pull apart.]

(b) RISK – People might get hurt or overheated if the room is too overcrowded, so this must be avoided.

(c) HOW IT MODELLED EQUILIBRIUM – People represented reactants and their joining up to dance represented the formation of products. Couples breaking up showed what happens when some products break up and form reactants again. Equilibrium was reached when the doors were closed (= a 'closed' system), as no-one could enter or leave the room and the number of couples dancing was constant. Although the number of people dancing stayed the same, it was not always the same couples– so the dancing was then in dynamic equilibrium – as it appeared to be staying the same. It was dynamic as there was constant change happening.

When the doors were opened for people to go in or out, a change in the concentration occurred and showed how it disturbed the equilibrium and that a new equilibrium position was eventually established with different numbers of people sitting and dancing. Closing off part of the room represented a decrease in volume and thus an increase in pressure – and led to more people becoming couples. This showed that the equilibrium had been disturbed, and that a new equilibrium was soon established.

LIMITATION – When people moved faster, they became tired and slowed down, whereas this does not happen with particles.

44. (a) $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

At time *A*: $[SO_2] = (0.8 - 0.3) = 0.5$ mol per 10 L $= 0.05$ mol L^{-1}

$[O_2] = (0.4 - 0.15) = 0.25$ mol per 10 L $= 0.025$ mol L^{-1}

$[SO_3] = (0.0 + 0.3) = 0.3$ mol per 10 L $= 0.03$ mol L^{-1}

$$K = \frac{[SO_3]^2}{[O_2][SO_2]^2} = \frac{[0.03]^2}{[0.025][0.05]^2} = 14.4$$

(b) Some time after time *A*, the temperature of the reaction vessel would have altered to favour the formation of products. At the new temperature, the value of *K* was higher and, after a time, a new equilibrium would have been established before time *B*.

[Note: The reaction shown is exothermic and so the temperature was decreased to favour products.]

45. (a) Initially – the system is at equilibrium. At 4 min, [products] begins to increase, [reactant] begins to decrease, so equilibrium is shifting to right. Since this reaction is endothermic ($\Delta H = +108$ kJ), the mixture must have been heated.

(b) At 10 minutes – some $CO(g)$ has been removed from the reaction mixture. By Le Châtelier's principle, the equilibrium will shift to the right to replace some of the CO that was removed. The [products] begins to increase, while [reactant] begins to decrease.

(c) At 14 min – $[Cl_2(g)]$, $[CO(g)]$ and $[COCl_2(g)]$ all decreased by 33%. This means a decrease in pressure occurred. There are 2 moles of gaseous products on the right and only 1 mole of gaseous reactants on the left. Le Châtelier's principle predicts that less pressure will shift the equilibrium towards the right to increase the total pressure to some extent in reaching a new equilibrium.

46. (a) $K = \dfrac{[I_2][Cl_2]^3}{[ICl_3]^2}$

(b) If $[Cl_2] = 0.45$ mol L^{-1}, then $[I_2] = 0.15$ mol L^{-1}.

So $[ICl_3] = 0.35 - 0.30 = 0.05$ mol L^{-1}

$$\therefore\ K = \frac{[I_2][Cl_2]^3}{[ICl_3]^2} = \frac{0.15 \times 0.45^3}{0.05^2} = 5.4675 \approx 5.47$$

(c) *Any TWO of the following:*

- The reaction will shift to the right producing more products.
- The equilibrium constant (*K*) will increase.
- The rate of the reaction will increase. [OR Equilibrium will be reached more quickly.]

47. (a) $K = \dfrac{[H_2O]^2}{[H_2S]^2[SO_2]}$

(b)

	Amt initially (mol L^{-1})	Amt at equilibrium (mol L^{-1})
H_2S	1.00	$1.00 - 0.5 = 0.5$
SO_2	1.00	$1.00 - 0.25 = 0.75$
H_2O	–	0.5

So:

$$K = \frac{[H_2O]^2}{[H_2S]^2[SO_2]} = \frac{0.5^2}{0.5^2 \times 0.75} = 1.33$$

(c) 1. Increased pressure.
2. Decreased temperature.
3. Increased concentration of SO_2.
4. Removal of water by adding dehydrating agent.

48. (a) $K = \dfrac{[CH_4(g)][H_2S(g)]^2}{[CS_2(g)][H_2(g)]^4}$

(b) The volume of the container is 5.0 L

$\therefore\ [CH_4(g)] = \frac{0.55}{5.0} = 0.11 \text{ mol L}^{-1}$ $\qquad [H_2S(g)] = \frac{0.125}{5.0} = 0.025 \text{ mol L}^{-1}$

$[CS_2(g)] = \frac{0.15}{5.0} = 0.030 \text{ mol L}^{-1}$ $\qquad [H_2(g)] = \frac{0.15}{5.0} = 0.030 \text{ mol L}^{-1}$

The equilibrium constant, $K = \dfrac{(0.11)(0.025)^2}{(0.030)(0.030)^4} = 2829 \approx 2800$ *(to 2 sig figs)*

49. (a)

Gas	*Quantity* (mol)	*Concentration* (mol L^{-1})
CH_4	2.00	$\frac{2.00}{5.0} = 0.40$
H_2O	1.25	$\frac{1.25}{5.0} = 0.25$
CO	0.75	$\frac{0.75}{5.0} = 0.15$
H_2	0.75	$\frac{0.75}{5.0} = 0.15$

$$K = \frac{[CO][H_2]^3}{[CH_4][H_2O]} = \frac{0.15 \times 0.15^3}{0.40 \times 0.25} = 5.06 \times 10^{-3}$$

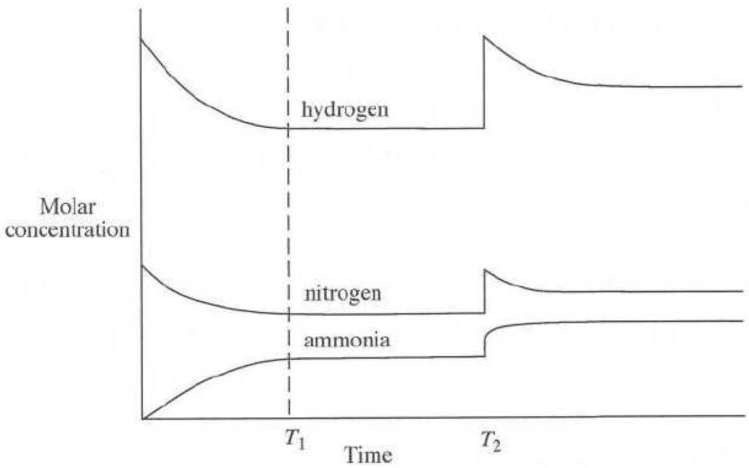
$\therefore\ K \neq 0.26$, so the reaction is not at equilibrium.

(b) The reaction is endothermic, so increasing the temperature will increase the yield. Increasing the concentration of one or both of the reactants will shift the equilibrium to the right, thus increasing the yield. The product side has 4 mol gas and the reactant side has 2 mol gas, so lowering the pressure will also shift the equilibrium to the right, thus increasing the yield.

[Note: Another adjustment would be to remove product as it is formed.]

50. (a) Equilibrium has been reached – the rate of the forward and reverse reactions has become equal.

(b) (i)



(ii) The increase in pressure increases the concentrations of both reactants and products. The equilibrium then shifts to partially counteract this change by moving to the side that has the smallest number of gaseous molecules, i.e. towards ammonia: $3H_2(g) + N_2(g) \rightleftharpoons 2NH_3(g)$

After a short time, a new equilibrium is established.

51. (a) Addition of catalyst – no effect on yield of N_2O_4 as a catalyst lowers activation energy for both directions, so the rates of both directions are increased equally. Increase in pressure – increases the yield of N_2O_4, as increased pressure moves the equilibrium in the direction which decreases the number of moles of gas.

(b)

	Initially	*At equilibrium*
NO_2	0.132 mol	Since 0.0800 mol reacted (using ratio in equation), there are 0.132 – 0.0800 = 0.052 mol at equilibrium.
N_2O_4	0 mol	0.0400 mol

Amounts at equilibrium are in mol L^{-1} as the vessel has 1 L volume.

At 25°C, $K = \frac{[N_2O_4]}{[NO_2]^2} = \frac{0.0400}{(0.052)^2} = 14.7929 \approx 14.8$

With the decrease in temperature from 100°C to 25°C, the K value has increased which indicates that the forward reaction is exothermic.

52. (a) This would have helped them to remove the outer layer and provided an increased surface area for the water to leach the toxins out more easily.

(b) Leaching involves soaking a solid substance in a liquid such as water, which allows substances that are soluble in the liquid to dissolve out of the solid into the liquid. The cycasin toxin is water soluble and so dissolves in water. Hence it can be removed from cycad seeds when they are soaked in water.

(c) Compared to sucrose, cycasin is relatively low in water solubility. So, to leach the toxin out of them, cycad seeds would need to be left in the running water over a period of time, as this would keep dissolving the toxin out at a maximum rate. However, when soaked in a wooden bowl, the seeds are in a fixed amount of water and so the toxin will only dissolve into the water until it reaches the equilibrium concentration. This risks some of the toxin remaining in the seeds and so using running water is preferable.

(d) The leaching method would have been developed as a result of learning what to do by trial and error to make the seeds safe to eat. So, it does not indicate that they had an understanding of the science related to solubility and equilibria.

53. (a) Colourimetry OR UV-visible spectroscopy

(b) $K_{eq} = \frac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^-]}$

(c) (i) increases

(ii) increases

(iii) decreases

(d) Adding water lowers the concentration of both the reactant and product species and so reduces the number of particles per volume. Hence there will be a shift in equilibrium towards the side with the most dissolved particles. Hence the reverse reaction will occur.

54. From the graph, the number of moles of each substance present in the 2 L container at equilibrium are:

$A = 4$ mol $\qquad B = 1$ mol $\qquad C = 3$ mol

$\therefore\ [A] = 2.0 \text{ mol L}^{-1} \qquad [B] = 0.5 \text{ mol L}^{-1} \qquad [C] = 1.5 \text{ mol L}^{-1}$

$\therefore$ equilibrium constant, $K = \frac{[C]^3}{[A][B]} = \frac{(1.5)^3}{(2.0)(0.5)} = 3.375 \approx 3.4$

55. (a) Light-brown → very pale to colourless

[Note: Pressure is decreased. In response to a decrease in pressure, more molecules are formed. As a result of equilibrium shifting to the left, the concentration of nitrogen dioxide which is brown decreases.]

(b) Light brown → a darker brown colour

[Note: More oxygen is added, and so equilibrium shifts to the right. As a result, more nitrogen dioxide molecules which are brown are produced.]

(c) No change in colour [Note: A catalyst does not change equilibrium position.]

56. From equation: $[Ag^+] = [Br^-]$

K_{sp} for this reaction $= [Ag^+][Br^-] = 5.0 \times 10^{-13} = [Ag^+]^2$

$\therefore [Ag^+] = [Br^-] = 7.1 \times 10^{-7}$ mol L^{-1}

57. (a) $Ca_2SO_4(s) \rightleftharpoons Ca^{2+}(aq) + SO_4^{2-}(aq)$

(b) $K_{sp} = [Ca^{2+}][SO_4^{2-}]$

(c) Solubility is 0.0045 mol L^{-1} at 25°C

$\therefore K_{sp} = (4.5 \times 10^{-3})(4.5 \times 10^{-3}) = 2.0 \times 10^{-5}$

(d) Yes – SO_4^{2-} ions from the sodium sulfate solution would be able to combine with Ca^{2+} that dissociate from $Ca_2SO_4(s)$. The SO_4^{2-} concentration in the solution will be higher than in water, so this will shift the equilibrium to the left. This will decrease the solubility of Ca_2SO_4.

58. (a) $Ag_2CrO_4 \rightleftharpoons 2Ag^+(aq) + CrO_4^{2-}(aq)$

Solubility of $Ag_2CrO_4 = 2.5 \times 10^{-2}$ g L^{-1}

Molar mass of $Ag_2CrO_4 = (107.9 \times 2) + (52.00 \times 1) + (16.00 \times 4) = 331.8$ g mol^{-1}

$\therefore$ moles of $Ag_2CrO_4 = \frac{2.5 \times 10^{-2}}{331.8} = 7.5 \times 10^{-5}$ mol L^{-1}

(b) $K_{sp} = [Ag^+]^2[CrO_4^{2-}] = (2 \times 7.5 \times 10^{-5})^2 \times (7.5 \times 10^{-5}) = 1.7 \times 10^{-12}$

(c) Let $[CrO_4^{2-}] = s$

$1.7 \times 10^{-12} = (0.1^2)\, s$

$\therefore s = \frac{1.7 \times 10^{-12}}{(0.1)^2}$

$= 1.7 \times 10^{-10}$ mol L^{-1}

(d) Dilute HNO_3 provides a higher $[H^+]$ than water.

$\therefore$ in dilute nitric acid:

$$2CrO_4^{2-}(aq) + 2H^+(aq) \rightleftharpoons 2HCrO_4^-(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + H_2O$$

... this will shift further to the right in dilute HNO_3 due to the higher $[H^+]$. Hence more silver chromate will dissolve in dilute HNO_3 to replace the CrO_4^{2-} ions removed.

59. (a) • *Solubility of AgCl*

$K_{sp} = [Ag^+][Cl^-] = (s)(s)$ ... using $[Ag^+] = s$

$2.0 \times 10^{-10} = (s)(s)$

$= s^2$

$\therefore s = 1.4 \times 10^{-5}$ mol L^{-1}

• *Solubility of* Ag_2CO_3

$K_{sp} = [Ag^+]^2[CO_3^{2-}] = (2s)^2(s)$ ... using $[CO_3^{2-}] = s$

$8.0 \times 10^{-12} = (4s^2)(s)$

$= 4s^3$

$\therefore s = 1.26 \times 10^{-4}$ mol L^{-1}

$\therefore$ Ag_2CO_3 is more soluble than AgCl.

(b) $K_{sp} = [Ag^+][Cl^-] = (s)(s + 0.1)$ ... using $[Ag^+] = s$

$2.0 \times 10^{-10} = (s)(s + 0.1)$

$= s^2 + 0.1s$ ... *as s << 0.1 the s^2 can be ignored*

$\therefore$ Solubility, $s = 2 \times 10^{-9}$ mol L^{-1}

60. As CO_2 is used up in photosynthesis, the equilibrium moves to the left as it attempts to release additional CO_2 into the water. In the process, additional OH^- is also formed. So the pH increases.

61. (a) Initial $[CO_2]$ = initial $[H_2]$ = 0.400 mol L^{-1}

mol CO formed = mol H_2O formed = mol H_2 reacted = mol CO_2 reacted

$[H_2] = 0.400 - 0.225 = 0.175$ mol L^{-1}

(b) $K = \dfrac{[CO(g)][H_2O(g)]}{[CO_2(g)][H_2(g)]}$

(c) $K = \dfrac{(0.225)\,(0.225)}{(0.400 - 0.225)(0.400 - 0.225)} = \dfrac{(0.225)\,(0.225)}{(0.175)(0.175)} = 1.65$

62. (a) No difference. OR The rates of the forward and reverse reactions are the same.

(b) This will reduce the pressure, so by Le Châtelier's Principle the equilibrium will shift to favour the reactants, as there are two moles of gaseous reactants and only one mole of gaseous products, i.e. more $PCl_3(g)$ and $Cl_2(g)$ will be produced.

(c) The equilibrium constant will remain unchanged – as it is only dependent on temperature and the temperature remains constant.

[Note: The equilibrium position will change towards the products if more $Cl_2(g)$ is added, NOT the equilibrium constant.]

63. (a) $K = \frac{[COCl_2]}{[CO][Cl_2]} = \frac{(0.4)}{(0.2)(0.1)} = 20$

(b) (i) CO

(ii) $[Cl_2]$ decreases

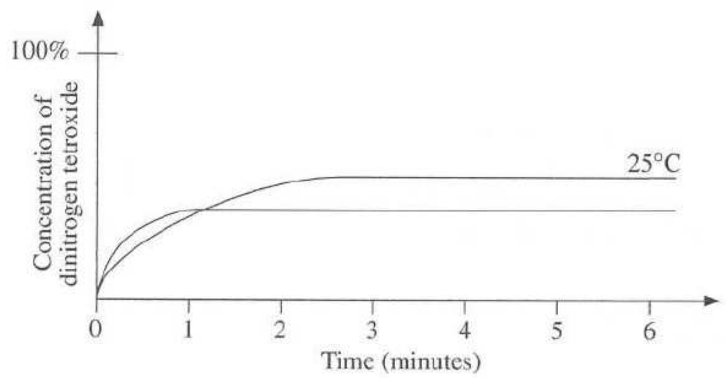
64. (a)

	$H_2(g)$	+	$I_2(g)$	$\rightleftharpoons$	$2HI(g)$
Mole ratio	1		1		2
$n_{initial}$	1 –		1 –		0 +
n_{end}	0.225		0.225		1.55
n_{change}	0.775		0.775		1.55
V	1		1		1

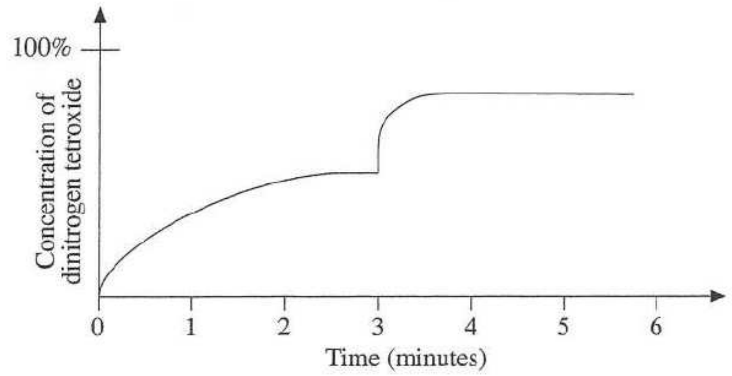
$\therefore$ Equilibrium concentration, $c = \frac{n}{v} = \frac{0.225}{1} = 0.225$ mol L^{-1} for both H_2 and I_2

(b) $K = \frac{[HI^2]}{[H_2][I_2]} = \frac{(1.55^2)}{(0.225)(0.225)} = 47.5$

65. (a)

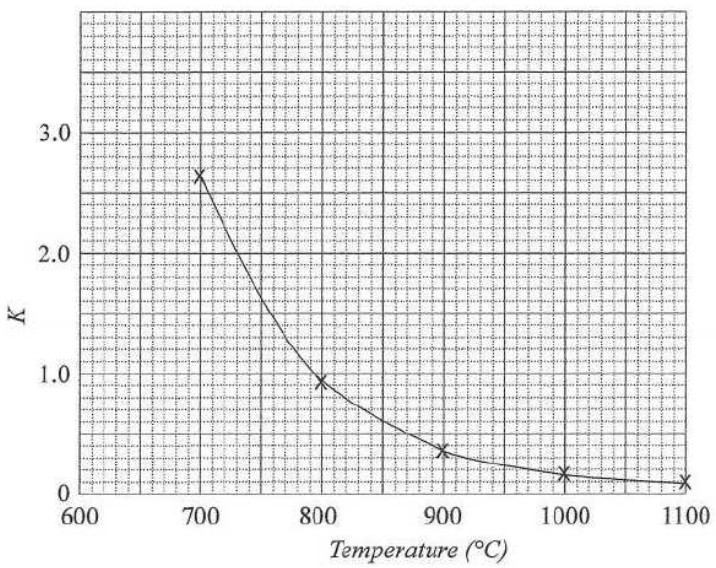


(b)



66. (a)

K versus temperature for SO_3 production



(b) As temperature increases, K decreases and so the [product] is decreasing, while the [reactant] is increasing. Increased temperature favours the reverse reaction by Le Châtelier's Principle. From the change in K, it can be seen that the reaction has shifted left – favouring the reactants being reformed, which is the endothermic reaction. So the forward reaction is exothermic, $\Delta H = -$ve.

67. (a) A reduction in the concentration of CO_2 would favour the reverse reaction. So more reactants would be formed.

(b) An increase in the temperature – since this is an exothermic reaction, this will shift the reaction to the left and more reactants will be formed.