



VCE CHEMISTRY 2009

YEAR 12 UNIT 4

Industrial Chemistry

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Time allowed: 50 minutes

Total marks: 47

12 Multiple Choice Questions

4 Short Answer Questions

An Answer Sheet is provided for Section A.

Answer all questions in Section B in the space provided.

To download the Chemistry Data Book please visit the VCAA website:

http://www.vcaa.vic.edu.au/vce/studies/chemistry/chem1_sample_2008.pdf Page 20

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Student Name.....

VCE Chemistry 2009 Year 12 Unit 4 - Industrial Chemistry

Student Answer Sheet

Instructions for completing test. Use only a 2B pencil. If you make a mistake erase and enter the correct answer. Marks will not be deducted for incorrect answers.

Write your answers to the Short Answer Section in the space provided directly below the question. There are 12 Multiple Choice questions to be answered by circling the correct letter in the table below.

Question 1 A B C D

Question 3 A B C D

Question 5 A B C D

Question 7 A B C D

Question 9 A B C D

Question 11 A B C D

Question 2 A B C D

Question 4 A B C D

Question 6 A B C D

Question 8 A B C D

Question 10 A B C D

Question 12 A B C D

VCE Chemistry 2009 Year 12 Unit 4 - Industrial Chemistry

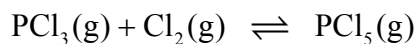
Multiple Choice Questions - Section A

(12 marks, 13 minutes)

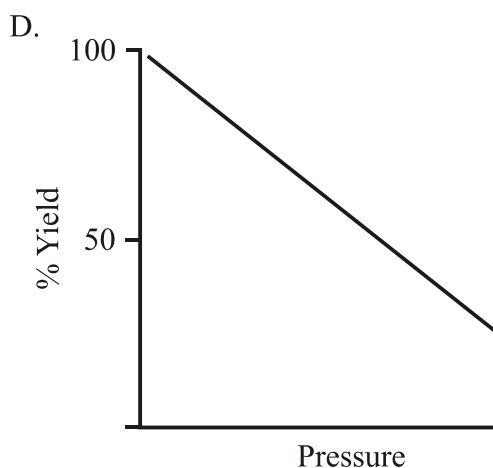
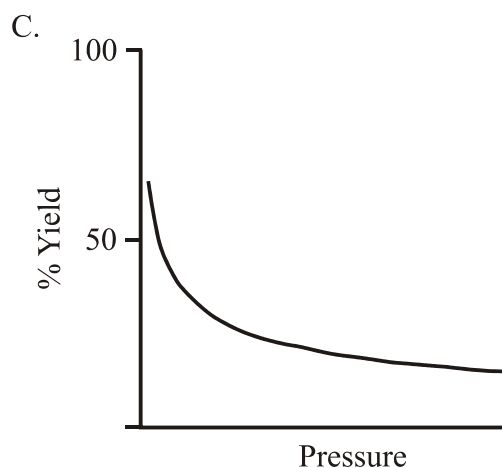
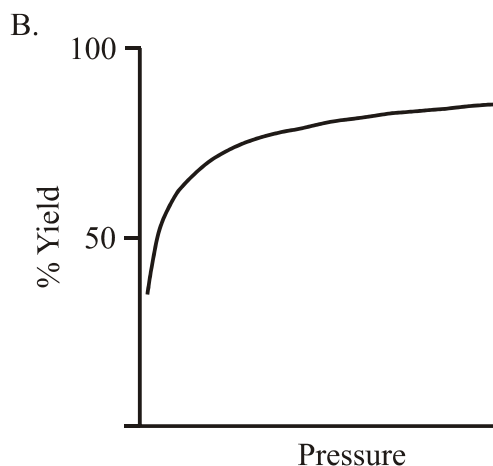
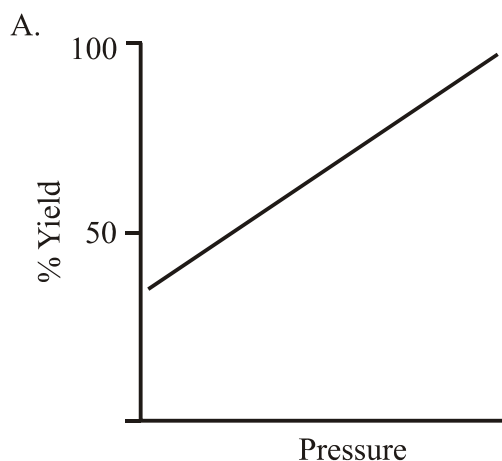
*This section contains 12 multiple choice questions.
For each question choose the response that is correct or best answers the question.
Indicate your answer on the answer sheet provided.
(Choose only **one** answer for each question.)*

Question 1

The reaction between phosphorous(III) chloride and chlorine can be represented by the chemical equation



Which one of the graphs below would best represent how the yield of phosphorous(V) chloride would be expected to behave as the pressure of the reaction is varied?



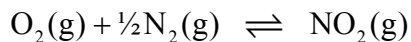
Question 2

The decomposition of nitrogen(IV) oxide can be represented by the chemical equation



The value for the equilibrium constant for this reaction at 600 K is 3.4×10^{10} M.

What would be the numerical value of the equilibrium constant for the reaction represented by the chemical equation at 600 K?



- A. 2.7×10^{-6}
- B. 1.5×10^{-11}
- C. 5.4×10^{-6}
- D. 2.9×10^{-11}

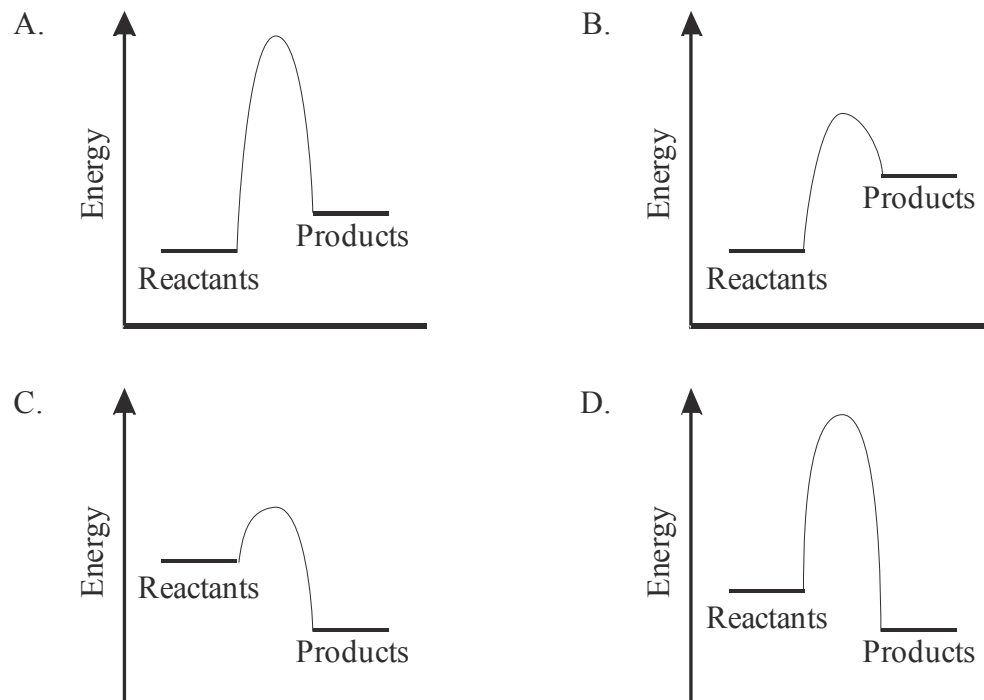
Question 3

The significant increase in rate of reaction for a chemical reaction that occurs when the temperature of the reaction is moderately increased, is best explained by

- A. the decreased rate of the reverse reaction at higher temperatures.
- B. the increased number of collisions between reactant particles with sufficient energy.
- C. the increased number of collisions between the reactant particles due to their increased speeds.
- D. the decrease in the activation energy at the higher temperature.

Question 4

The rate of a chemical reaction is slow in the absence of a catalyst and the equilibrium yield decreases when the temperature is increased. The energy profile that would best represent this reaction would be



Question 5

The addition of some solid sodium propanoate to a 0.10 M aqueous solution of propanoic acid resulted in the equilibrium concentration of the propanoate ion being 5.2×10^{-3} M. The pH of this solution would be

- A. 3.6
- B. 2.9
- C. 2.3
- D. 6.2

Question 6

The activation energy can best be described as

- A. the minimum amount of energy that the reactant particles must have for a fruitful reaction to occur.
- B. the energy required to break all of the bonds in the reactant particles so that the reaction proceeds.
- C. the change in energy that occurs as a reaction proceeds to equilibrium.
- D. the maximum amount of energy that is either absorbed or released by the reaction before it reaches equilibrium.

Question 7

The pharmaceutical industry tends to produce much of its products using batch processing methods. Which one of the following would **not** be a reason why the pharmaceutical industry uses batch processing?

- A. The amounts of materials produced are relatively small.
- B. The reactants often have to be added in specific amounts to produce the desired products.
- C. Batch processing methods use less expensive equipment and cheap energy sources.
- D. Many of the starting materials are expensive and/or legally controlled substances that must be accounted for.

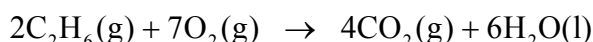
Question 8

An aqueous solution was prepared by dissolving 0.0200 g of sodium hydroxide in 500.0 mL of deionised water. The pH of this solution would be closest to

- A. 11.0
- B. 3.0
- C. 3.3
- D. 10.7

Question 9

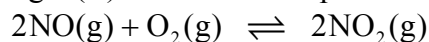
The change in enthalpy, ΔH , for the reaction described by the chemical equation below would be



- A. $+3114 \text{ kJ mol}^{-1}$
- B. $-3114 \text{ kJ mol}^{-1}$
- C. $-1557 \text{ kJ mol}^{-1}$
- D. $+1557 \text{ kJ mol}^{-1}$

Question 10

The gas phase oxidation of nitrogen(II) oxide can be represented by the chemical equation



The value of the equilibrium constant for this reaction at 200 °C is $3 \times 10^6 \text{ M}^{-1}$. This information suggests that

- A. the forward reaction is endothermic.
- B. at equilibrium the position of equilibrium will favour the reactants.
- C. the forward reaction is exothermic.
- D. at equilibrium the position of equilibrium will favour the products.

Question 11

The reaction used in an industrial process to produce hydrocarbons can be described by the chemical equation



The following changes can be made to the industrial process:

- I Increase the temperature.
- II Increase the pressure.
- III Use a catalyst.
- IV Increase the amount of hydrogen in the feed gas.

Which of these changes would improve the equilibrium yield of the hydrocarbon?

- A. I and II.
- B. I and III.
- C. II and IV.
- D. II, III and IV.

Question 12

In industrial waste management practices, which one of the following strategies is considered as the best practice?

- A. Treatment of waste.
- B. Reduction of waste materials formed.
- C. Recycling of waste materials formed.
- D. Responsible disposal of waste materials.

End of Section A

VCE Chemistry 2009 Year 12 Unit 4 - Industrial Chemistry

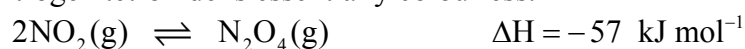
Short Answer Questions – Section B

(35 marks, 37 minutes)

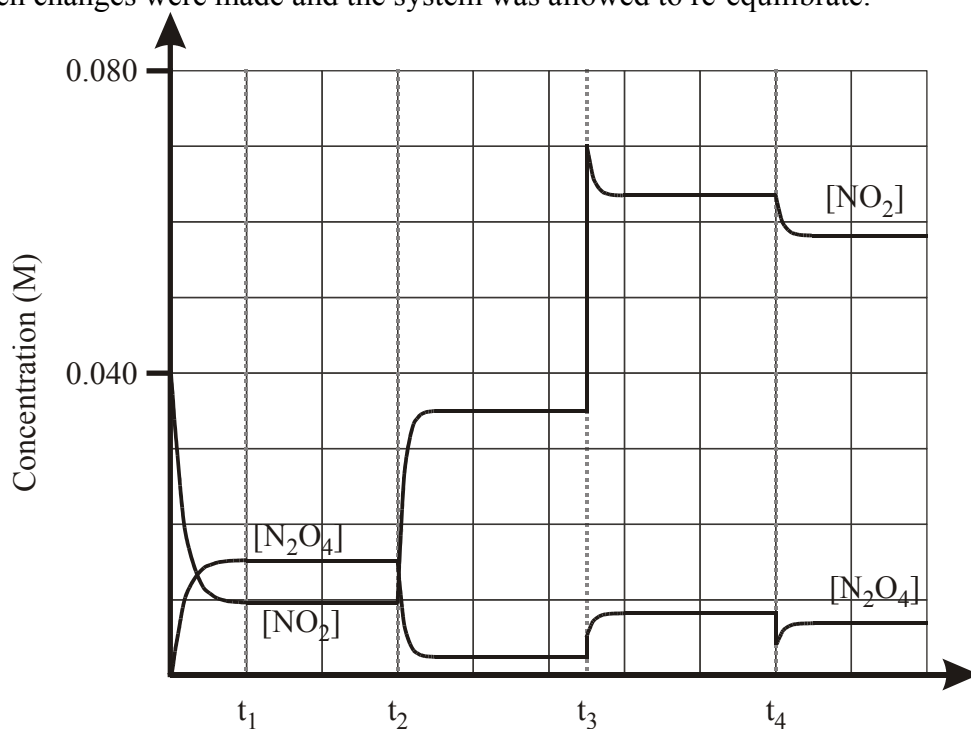
*This section contains four questions, numbered 1 to 4.
All questions should be answered in the spaces provided.
The mark allocation and approximate time that should be spent on each question are given.*

Question 1 (10 marks, 11 minutes)

Nitrogen dioxide, NO_2 , readily undergoes a dimerisation reaction to form dinitrogen tetroxides, N_2O_4 . This reaction is an easy one to study because the nitrogen dioxide is dark brown and the dinitrogen tetroxide is essentially colourless.



A pure sample of nitrogen dioxide was placed in a container and allowed to equilibrate at 25°C . The graph below also shows how the concentrations of the two gases changed with time when changes were made and the system was allowed to re-equilibrate.



- a. Write an expression for the equilibrium constant for this reaction.

(1 mark)

- b. From the data in the graph, determine the value of the equilibrium constant after it had reached equilibrium at 25 °C, just following t_1 .

(1 mark)

- c. At t_2 a single change was made to the system's conditions before it was allowed to re-equilibrate at the new set of conditions.

- i. What condition was changed at this time?
- ii. Explain why this change resulted in the altered concentrations of the two gases when equilibrium has been re-established.
- iii. How would this change in conditions affect the value for the equilibrium constant?

(1 + 1 + 1 = 3 marks)

- d. A further change was made to the system's conditions at t_3 before the system was allowed to re-equilibrate at the new set of conditions.

- i. What condition was changed at this time?
- ii. Explain why this change resulted in the altered concentrations of the two gases when equilibrium has been re-established.

- iii. How would this change in conditions affect the value for the equilibrium constant?

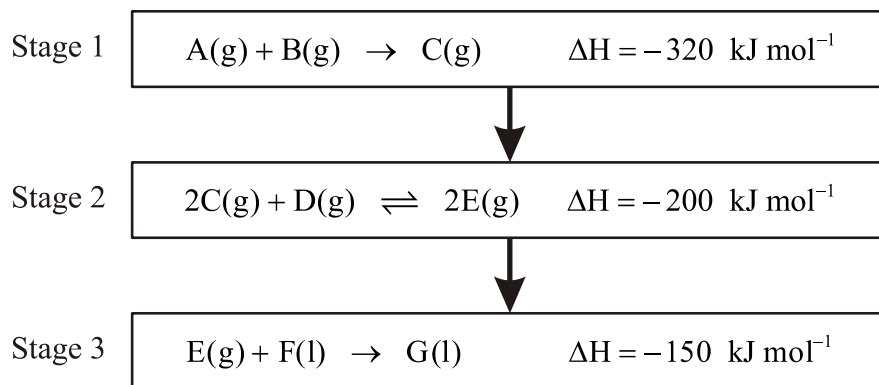
(1 + 1 + 1 = 3 marks)

- e. A final change was made to the system's conditions at t_4 before it was allowed to re-equilibrate at the new set of conditions.
- i. What condition was changed at this time?
- ii. Explain why this change resulted in the altered concentrations of the two gases when equilibrium has been re-established.

(1 + 1 = 2 marks)

Question 2 (8 marks, 8 minutes)

Three stages of an industrial process can be represented by the following flow chart.



- a. The value of the equilibrium constant for the second stage reaction is large, however the reaction rate is slow at 100 °C. What are two methods that could be used to improve the reaction rate for this stage of the process?
- i.
- ii.

(1 mark)

- b. Explain how each of the changes proposed in a. leads to an increased reaction rate for this reaction?

i.

ii.

(1 + 1 = 2 marks)

- c. Explain the effect of each of the changes proposed in a. above on the equilibrium yield for this reaction?

i.

ii.

(1 + 1 = 2 marks)

- d. What change, other than those proposed in a. above, could be made to improve the equilibrium yield for the second stage reaction?

(1 mark)

- e. How could the energy change that results from the reaction in the first stage of the process be used within the chemical plant?

(1 mark)

- f. Under the conditions that the plant managers use to best maximise equilibrium and rate considerations, the gas stream exiting the second stage contains a mixture of all three gases. Gas C is separated from this gas stream before it enters the third stage of the process. Gas C is toxic to the environment. What would be the best method that the chemical plant managers could use to deal with this gas?

(1 mark)

Question 3 (10 marks, 11 minutes)

- a. The ammonium ion is a very weak acid.

- i. Write an appropriate chemical equation to show the ammonium ion acting as an acid in an aqueous solution.

(1 mark)

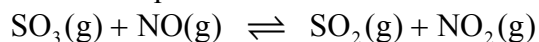
- ii. Write an expression for the acidity constant for the ammonium ion.

(1 mark)

- iii. Calculate the pH of an aqueous solution of ammonium sulfate, $(\text{NH}_4)_2\text{SO}_4$, prepared by dissolving 3.50 g of solid in 250.0 mL of deionised water.

(4 marks)

- b. The reaction between sulfur(VI) oxide, SO_3 , and nitrogen(II) oxide, NO , can be represented by the chemical equation



0.080 mol each of sulfur(VI) oxide and nitrogen(II) oxide were mixed in 2.0 L vessel at 800 °C and allowed to reach equilibrium. Analysis of the equilibrium mixture indicated that the concentration of nitrogen(IV) oxide, NO_2 , was 0.0090 M.

- i. Write an expression for the equilibrium constant.

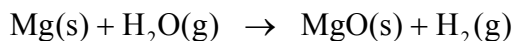
(1 mark)

- ii. Determine the value for the equilibrium constant for this reaction at 800 °C.

(3 marks)

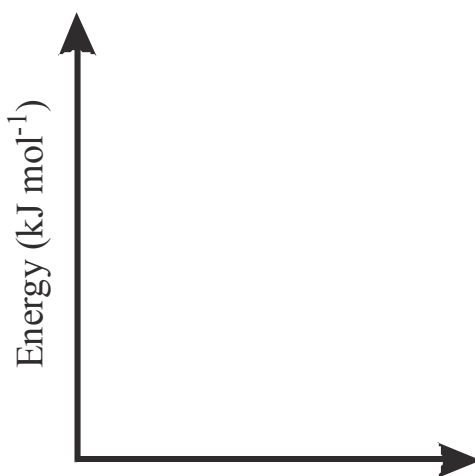
Question 4 (7 marks, 7 minutes)

- a. A reaction between magnesium metal and water vapour can be described by the chemical equation



The chemical potential energy of the reactants and products for this reaction are -242 kJ mol^{-1} and -602 kJ mol^{-1} respectively. The activation energy for the forward reaction is 926 kJ mol^{-1} .

- i. On the grid below sketch the energy profile for this reaction.



(1 mark)

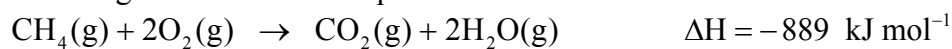
ii. Determine the enthalpy change for this reaction.

(1 mark)

iii. What is the activation energy for the reverse reaction?

(1 mark)

b. The complete oxidation of methane and carbon monoxide can be described by the following thermochemical equations



i. Write an appropriate chemical equation for the incomplete oxidation of methane to form carbon monoxide and water vapour.

(1 mark)

ii. Use the thermochemical equations to determine the enthalpy change, ΔH , for the reaction described by the chemical equation written in i.

(2 marks)

c. Explain why the activation energy for any chemical reaction must be a positive quantity.

(1 mark)

End of Section B

End of Unit Test

Suggested Answers

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Multiple Choice Questions – Section A

(1 mark per question)

- Q1 B In this reaction the **number of mole of products is less than the number of moles of reactants**, therefore using **Le Chatelier's Principal** an **increase in pressure will shift the position of equilibrium to favour the side of the reaction with the lesser number of moles**. In this case it is the formation of phosphorous(V) chloride, therefore **the yield would be expected to increase as the pressure increases**. Response A shows a linear relationship which if extrapolated further would result in a yield greater than 100 %, which cannot occur, therefore it is not an acceptable answer.

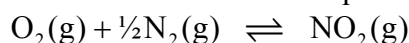
- Q2 C The expression for the equilibrium constant for the reaction
 $2\text{NO}_2(\text{g}) \rightleftharpoons \text{N}_2(\text{g}) + 2\text{O}_2(\text{g})$

$$\text{is } K_1 = \frac{[\text{N}_2][\text{O}_2]^2}{[\text{NO}_2]^2} = 3.4 \times 10^{10}$$

The value for the equilibrium constant for any reverse reaction, K_r , is the reciprocal of the value of the forward reaction equilibrium constant, K_f .

$$K_r = \frac{1}{K_f}$$

For the reverse reaction represented by the chemical equation



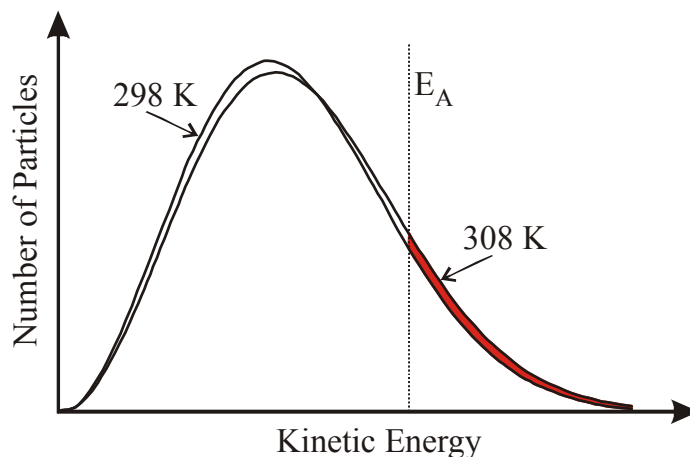
When the stoichiometry for the reaction is halved, the value for the new equilibrium constant, K' , will be square root of the value of the equilibrium constant.

$$K' = \sqrt{K}$$

$$K_2 = \frac{[\text{NO}_2]}{[\text{O}_2][\text{N}_2]^{1/2}} = \sqrt{\frac{1}{K_1}} = \sqrt{\frac{1}{3.4 \times 10^{10}}} = 5.4 \times 10^{-6} \text{ M}^{-1/2}.$$

- Q3 B An increase in temperature of a reaction will increase the kinetic energy of the reactants, this has two effects. Firstly the particles will be travelling faster, therefore there is a higher chance of collision between the particles, but this alone cannot explain the significant increases in rates. The second more important factor is that more particles will have sufficient energy to overcome the activation energy, E_A , 'barrier' and result in the occurrence of a fruitful reaction.

The diagram on the right shows the increased number of gas particles with energy greater than the activation energy following a 10 K temperature increase.



- Q4 D Since the rate of the uncatalysed reaction is slow, this implies that the **activation energy, E_A , for the reaction is a significant factor.**
The reduction in the equilibrium yield with increasing temperature implies that the reaction is **exothermic, $\Delta H < 0$.**
The energy profile diagram that best supports these criteria is that in response **D.**
- Q5 A Propanoic acid, $\text{CH}_3\text{CH}_2\text{COOH}$, $\text{C}_2\text{H}_5\text{COOH}$, is a weak acid and in a 0.10 M aqueous solution the $[\text{C}_2\text{H}_5\text{COOH}]$ would be close to 0.10 M.
From Table 12 of the *VCE Chemistry Data Booklet* the acidity constant, K_a , is 1.3×10^{-5} M.
$$\text{C}_2\text{H}_5\text{COOH}(\text{aq}) \rightleftharpoons \text{C}_2\text{H}_5\text{COO}^-(\text{aq}) + \text{H}^+(\text{aq})$$

The equilibrium concentration of the propanoate ion after the solid sodium propanoate had been added $[\text{C}_2\text{H}_5\text{COO}^-] = 5.2 \times 10^{-3}$ M.
The expression for the acidity constant is
$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_5\text{COO}^-]}{[\text{C}_2\text{H}_5\text{COOH}]} = 1.3 \times 10^{-5}$$

Substituting the values into this expression
$$1.3 \times 10^{-5} = \frac{[\text{H}^+](5.2 \times 10^{-3})}{(0.10)}$$

$$[\text{H}^+] = \frac{(1.3 \times 10^{-5})(0.10)}{(5.2 \times 10^{-3})} = 2.5 \times 10^{-4} \text{ M}$$

$$\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(2.5 \times 10^{-4}) = 3.6$$
- Q6 A The activation energy is the **minimum amount of energy** that the reactant particles require for a reaction to proceed and be fruitful. This is the energy required to overcome the ‘energy barrier’ for the reaction.
- Q7 C Batch processing methods are best suited for the production of relatively small amounts of materials, where the cost and delays that result from restarting the process are not a major factor in the overall cost. Batch processing is also the better approach when the amounts of materials being mixed needs to be closely controlled either for a reaction stoichiometry, cost and/or accountability reasons. Batch processing equipment tends to be more expensive and designed to carry out specific processes, than the equipment used in continuous flow processes.
- Q8 A $M(\text{NaOH}) = 23.0 + 16.0 + 1.0 = 40.0 \text{ g mol}^{-1}$
$$\text{NaOH}(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$$

$$n(\text{NaOH}) = m / M = 0.0200 / 40.0 = 5.00 \times 10^{-4} \text{ mol}$$

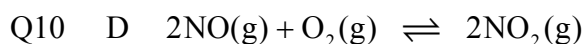
$$n(\text{OH}^-) = n(\text{NaOH}) = 5.00 \times 10^{-4} \text{ mol}$$

$$[\text{OH}^-] = n / V = 5.00 \times 10^{-4} / (500/1000) = 1.00 \times 10^{-3} \text{ M}$$

From Table 3 *VCE Chemistry Data Booklet*, $K_w = 1.00 \times 10^{-14} \text{ M}^2$
$$K_w = [\text{H}^+][\text{OH}^-] = 1.00 \times 10^{-14} = [\text{H}^+](1.00 \times 10^{-3})$$

$$[\text{H}^+] = 1.00 \times 10^{-14} / 1.0 \times 10^{-3} = 1.00 \times 10^{-11} \text{ M}$$

$$\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(1.00 \times 10^{-11}) = 11.0$$
- Q9 B From **Table 13** the combustion of one mole of ethane releases 1557 kJ energy, therefore since energy has been released **this is an exothermic reaction and $\Delta H < 0$.**
The equation has 2 mol of ethane reacting therefore
$$\Delta H = 2 \times (-1557) = \mathbf{3114 \text{ kJ mol}^{-1}}.$$



The **value of the equilibrium constant is much greater than 1**, therefore at equilibrium, the **position of equilibrium will favour the formation of the product** over the reactants. The value of the equilibrium constant at a single temperature cannot be used to determine if a reaction is endothermic or exothermic.

When a stoichiometric mixture of these reactants is allowed to reach equilibrium, the expected yield of nitrogen(IV) oxide, based on the value of the equilibrium constant, would be greater than 90 %.

Q11 C I The reaction is **exothermic**, therefore an increase in temperature would result in a **decrease in the equilibrium yield**, and the value for the equilibrium constant would decrease.

II Increasing the pressure, using Le Chatelier's Principal, **will favour the side of the reaction with the lesser number of mole**. In this case it will favour the products and **increase the equilibrium yield**.

III The addition of a **catalyst has no effect on the equilibrium yield** only on the speed that equilibrium is reached.

IV Increasing the amount of hydrogen in the feed stock will, using Le Chatelier's Principal, **shift the position of equilibrium to the right, increasing the amount of products formed and the equilibrium yield**.

Changes II and IV produce the desired effect.

Q12 B **Reduction in the amount of waste produced** is the best strategy of those listed. Treatment and disposal involve extra costs as does recycling and the waste is still produced. Other strategies that can be employed, and not provided as alternatives, are better and involve the prevention of waste formation and the elimination of waste in the process itself.

Short Answer (Answers) – Section B

Question 1 (10 marks, 11 minutes)



a. $K = \frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2}$ (1 mark)

b. From the graph the equilibrium concentrations are:

$[\text{NO}_2] = 0.010 \text{ M}$

$[\text{N}_2\text{O}_4] = 0.015 \text{ M}$

$K = \frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2} = \frac{(0.015)}{(0.010)^2} = 150 \text{ M}^{-1}$. (1 mark)

- c. i. At t_2 the concentration of the NO_2 increases while the concentration of N_2O_4 decreases. The only change that could cause this is a temperature change. Since the forward reaction is exothermic, then **an increase in temperature** will favour the reverse reaction. (1 mark)
- ii. An increase in **temperature will favour the reverse reaction** therefore the **position of equilibrium will shift to the left** and the **NO_2 concentration will increase while the N_2O_4 concentration will decrease**. (1 mark)

- iii. **Temperature is the only condition change that can change the value of the equilibrium constant.** Since the **position of equilibrium has shifted to the left, then the value of the equilibrium constant will have decreased.** (1 mark)
- d.
 - i. Both concentrations have increased, therefore either the **pressure has increased** or the **volume occupied by the gases has decreased.** (1 mark)
 - ii. Using Le Chatelier's Principal that the system will adjust its position to oppose the effect of the change, **an increase in pressure will favour the side of the reaction with fewer particles.** Therefore the position of equilibrium for this reaction will **shift to favour the formation of N₂O₄.** (1 mark)
 - iii. The position of equilibrium has shifted but the **value of the equilibrium constant will not be changed** by this change in conditions. (1 mark)
- e.
 - i. At t₄ only the **concentration of the N₂O₄ drops**, suggesting the **removal of some of this material.** (1 mark)
 - ii. The removal of some of the product will result in the position of equilibrium shifting to oppose this change and the **forward reaction will be favoured until equilibrium is re-established.** (1 mark) The [N₂O₄] increases from its value following the change at t₄, because the forward reaction is proceeding, but is less than its value prior to the change when equilibrium is re-established.

Question 2 (8 marks, 8 minutes)

- a. Second stage reaction: $2\text{C}(\text{g}) + \text{D}(\text{g}) \rightleftharpoons 2\text{E}(\text{g}) \quad \Delta H = -200 \text{ kJ mol}^{-1}$
 The thermochemical nature of the reaction has no bearing on the methods chosen to increase the rate of reaction.
Possible answers any two. (1 mark)
Increase the temperature.
Increase the pressure of the system, thereby increasing the concentration of the reactants.
Add a catalyst.
- b. **(Correct answer for change chosen 1 mark. Total marks = 2 marks)**
Increase the temperature
 An increase in temperature will result in **more particles having a higher kinetic energy and an increased chance of a fruitful reaction occurring** when a collision takes place.
Increase the pressure of the system
 Increasing the pressure of the system increases the number of particles occupying a volume, therefore there is a **higher likelihood of a fruitful collision occurring.**
Addition of a catalyst
 The addition of a catalyst **provides an alternative pathway with a lower activation energy**, therefore for a given temperature more particles will have this lower energy, increasing the likelihood of a fruitful collision occurring.
- c. **(Correct answer for change chosen 1 mark. Total marks = 2 marks)**
Increase the temperature
 The reaction in Stage 2 is exothermic, therefore an increase in temperature will **favour the reverse reaction and shift the position of equilibrium to the left, thereby lowering the equilibrium yield.**

Increase the pressure of the system

An increase in pressure will **favour the side with fewer particles**. The reaction in Stage 2 has three particles on the left and two on the right, therefore the **forward reaction would be favoured and the equilibrium position will shift to the right**. This will **increase the equilibrium yield**.

Addition of a catalyst

This has **no effect** on the position of equilibrium only on the speed that the equilibrium is reached.

- d. Possible choices include: **(1 mark)**

Increasing the pressure at which the reaction is carried out. **(If not chosen as in a.)**

Increasing the concentration of one of the reactants, such as D, in the feed mixture.

Removal of some of the product E from the gas mixture.

- e. The first stage reaction: $A(g) + B(g) \rightarrow C(g)$ $\Delta H = -320 \text{ kJ mol}^{-1}$

This reaction is exothermic and the heat from this reaction could be used:

to **heat the gases in the second stage**, or

to **heat water to produce steam** that could be used in another part of the plant, or

to **produce steam** that could be used to **drive a turbine to produce electricity** for use in the plant or added to the electricity grid.

(1 mark for any of the alternatives)

- f. The best practice would be to **recycle gas C back into the feed for the second stage**. **(1 mark)**

Question 3 (10 marks, 11 minutes)

- a. i. An acid is a proton donor.



- ii. $K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} \quad \text{(1 mark)}$

- iii. $M((\text{NH}_4)_2\text{SO}_4) = 14.0 \times 2 + 1.0 \times 8 + 32.1 + 16.0 \times 4 = 132.1 \text{ g mol}^{-1}$

$$n((\text{NH}_4)_2\text{SO}_4) = m / M = 3.50 / 132.1 = 2.65 \times 10^{-2} \text{ mol}$$

$$c((\text{NH}_4)_2\text{SO}_4) = n / V = 2.65 \times 10^{-2} / (250.0/1000) = \mathbf{1.06 \times 10^{-1} \text{ M}} \quad \text{(1 mark)}$$

$$[\text{NH}_4^+] = 2 \times c((\text{NH}_4)_2\text{SO}_4) = 2 \times 1.06 \times 10^{-1} = \mathbf{2.12 \times 10^{-1} \text{ M}} \quad \text{(1 mark)}$$

In this equilibrium reaction $[\text{NH}_3] = [\text{H}_3\text{O}^+] = x \text{ M}$

Since the ammonium ion is a weak acid then the concentration of the ammonium ion will be essentially unchanged at equilibrium.

The acidity constant for the ammonium ion can be found in Table 12

$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = \frac{x^2}{2.12 \times 10^{-1}} = 5.6 \times 10^{-10}$$

$$x^2 = 1.2 \times 10^{-10}$$

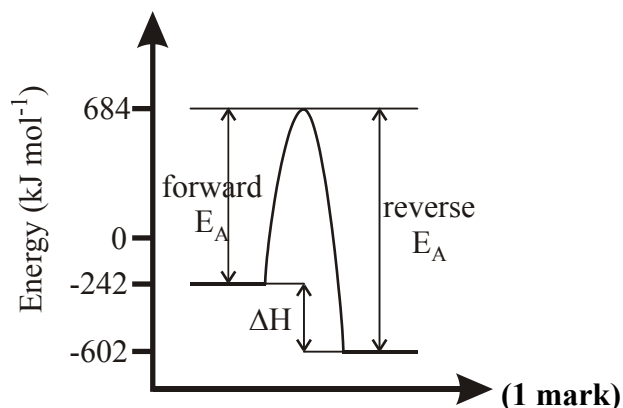
$$[\text{H}_3\text{O}^+] = \mathbf{1.1 \times 10^{-5} \text{ M}} \quad \text{(1 mark)}$$

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log(1.1 \times 10^{-5}) = \mathbf{5.0} \quad \text{(1 mark)}$$

- b. i. $K = \frac{[\text{SO}_2][\text{NO}_2]}{[\text{SO}_3][\text{NO}]}$ (1 mark)
- ii. At equilibrium $[\text{SO}_2] = [\text{NO}_2] = 0.0090 \text{ M}$
 The initial concentrations of SO_3 and NO were equal
 $[\text{SO}_3]_i = [\text{NO}]_i = n / V = 0.080 / 2.0 = 0.040 \text{ M}$ (1 mark)
 For every mole of NO_2 formed 1 mole of SO_3 reacted with 1 mole of NO .
 Therefore at equilibrium:
 $[\text{NO}]_e = 0.040 - 0.0090 = 0.031 \text{ M} = [\text{SO}_3]_e$ (1 mark)
 $K = \frac{(0.0090)(0.0090)}{(0.031)(0.031)} = 8.4 \times 10^{-2}$ (1 mark)

Question 4 (7 marks, 7 minutes)

a. i.

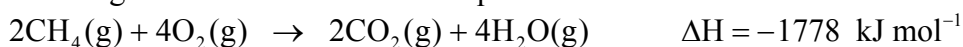


- ii. The enthalpy change is the difference in the chemical potential energies of the products and reactants.
 $\Delta H = -602 - (-242) = -360 \text{ kJ mol}^{-1}$ (1 mark)
- iii. From the graph the activation energy for the reverse reaction will be;
 $E_A = 684 + 602 = 1286 \text{ kJ mol}^{-1}$ (1 mark)
 Alternatively;
 The activation energy for the forward reaction is 926 kJ mol^{-1} .
 The chemical potential energy of the products is 360 kJ mol^{-1} less than that of the reactants, therefore the activation energy for the reverse reaction will be
 $E_A(\text{reverse reaction}) = 926 + 360 = 1286 \text{ kJ mol}^{-1}$. (1 mark)
- b. i. The products of the reaction will be CO and H_2O , therefore the unbalanced chemical equation will be
 $\text{CH}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}(\text{g}) + \text{H}_2\text{O}(\text{g})$
 Balance the oxygen atoms
 $\text{CH}_4(\text{g}) + \frac{3}{2}\text{O}_2(\text{g}) \rightarrow \text{CO}(\text{g}) + 2\text{H}_2\text{O}(\text{g})$
 Multiplying the coefficients by 2 will result in
 $2\text{CH}_4(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}(\text{g}) + 4\text{H}_2\text{O}(\text{g})$ (1 mark)

- ii. The thermochemical equations are:



Doubling the first thermochemical equation will result in



Reversing the second thermochemical equation will result in



(1 mark)

Adding these two thermochemical equations will result in the desired reaction and the enthalpy change will be the sum of the two enthalpy changes.

$$\Delta H = -1778 + 566 = -1212 \text{ kJ mol}^{-1}$$



(1 mark)

- c. For a chemical reaction to proceed, additional energy is needed to break the required bonds in the reacting particles. The particles must have this additional energy for a fruitful reaction to occur, therefore this quantity will always be positive. **(1 mark)**
This is clearly seen in the energy profile shown above.

End of Suggested Answers