# RATES AND EQUILIBRIUM

## PAST EXAM QUESTIONS

**WACE 2016 Sample Q1:** *Originally from WACE 2012*

The reaction of iron(III) oxide with carbon monoxide gas is shown below:

$$Fe_2O_3(s) + 3\,CO(g) \rightleftharpoons 2\,Fe(\ell) + 3\,CO_2(g)$$

Which one of the following changes to the system will initially decrease the rate of the forward reaction?

(a) decreasing the volume of the reaction vessel
(b) decreasing the pressure of CO(g) in the vessel
(c) decreasing the $Fe_2O_3$(s) particle size
(d) decreasing the concentration of $CO_2$(g) in the system

**WACE 2012 Q18:**

Ammonium chloride ($NH_4C\ell$) dissolves readily in water at room temperature. If a sample of ammonium chloride is dissolved in a beaker of water, the beaker becomes cold to the touch. Which one of the following is the **best** explanation for this observation?

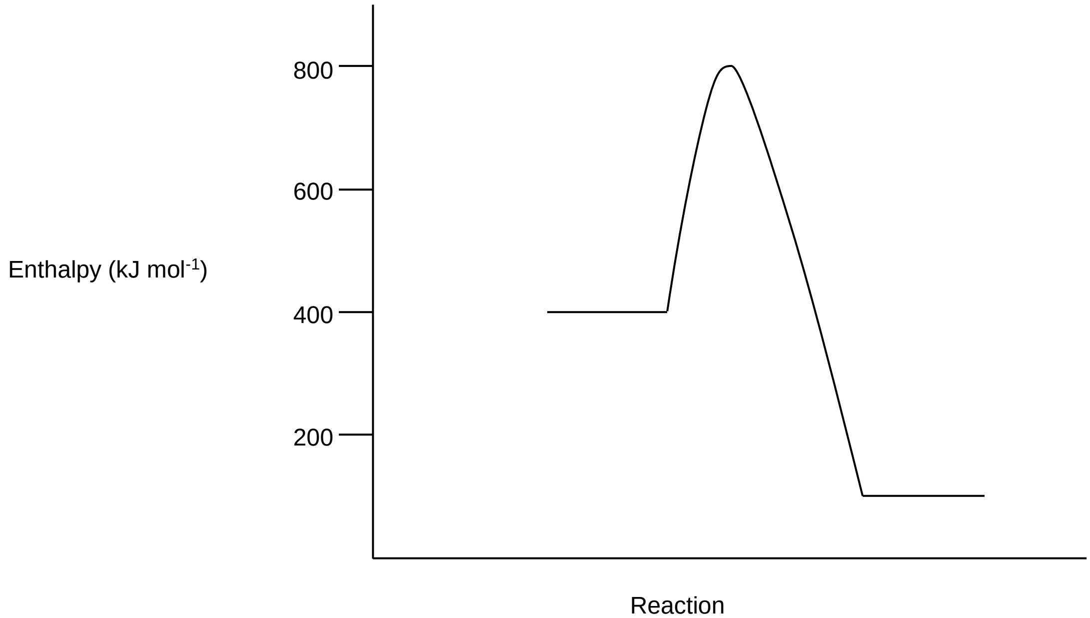
(a) The reaction is exothermic with a small activation energy
(b) The reaction is exothermic with a large activation energy
(c) The reaction is endothermic with a small activation energy
(d) The reaction is endothermic with a large activation energy

**WACE 2010 Q17:**

A small rise in temperature of gaseous reactants in a system results in an increase in the rate of reaction. Which one of the following is the **main** reason for this change?

(a) an increase in the speed of reactant particles, leading to a higher rate of collision
(b) an increase in the pressure inside the reaction vessel, leading to a higher rate of collision
(c) an increase in the proportion of collisions with more than the activation energy
(d) an increase in the activation energy of the reaction

*Use the potential energy diagram shown below to answer TEE 2009 questions 11 and 12.*




**TEE 2009 MC Q11:**

Which one of the following gives the correct values for the enthalpy change (ΔH) and the activation energy ($E_a$) for the forward reaction?

| | ΔH (kJ mol$^{-1}$) | $E_a$ (kJ mol$^{-1}$) |
|---|---|---|
| (a) | -300 | +700 |
| (b) | +300 | +400 |
| (c) | -300 | +800 |
| (d) | -300 | +400 |

**TEE 2009 MC Q12:**

A catalyst was added to the reaction mixture. Comparing the catalysed reaction with the uncatalysed reaction, which one of the following will remain the same?

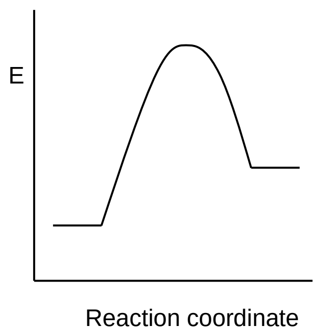
(a) the enthalpy change of the reaction
(b) the activation energy of the forward reaction
(c) the energy of the transition state
(d) the activation energy for the reverse reaction

**TEE 2007 MC Q9:**

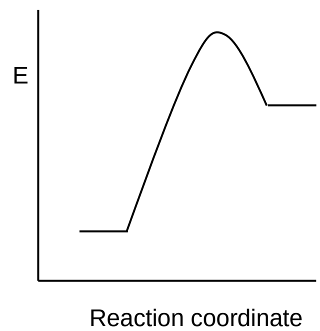
Nitroglycerine is a highly dangerous explosive substance. Simply dropping a container of nitroglycerine provides enough kinetic energy on impact with the floor to cause it to explode, releasing a very large amount of energy.

Which of the following energy profile diagrams would **best** represent this reaction?

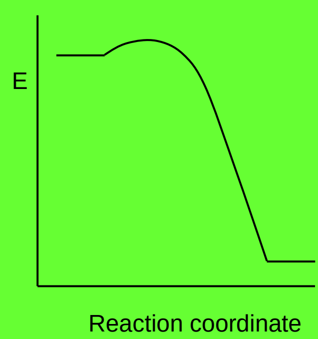
(a)



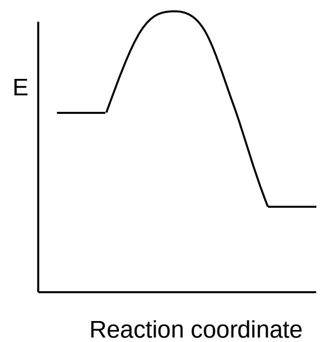

(b)




(c)




(d)




**TEE 2006 MC Q25:**

Finely ground aluminium and iron(III) oxide powders are mixed and placed in a container. As this reaction does not occur at room temperature, a burning piece of magnesium is dropped onto the mixture to ignite it. A bright, hot flame is observed. Which one of the following statements about the reaction is true?

(a) The magnesium acts as a catalyst and the reaction is exothermic.
(b) The magnesium provides the activation energy and the reaction is exothermic.
(c) The magnesium acts as a catalyst and the reaction is endothermic.
(d) The magnesium provides the activation energy and the reaction is endothermic.

*Use the following information to answer TEE 2000 Questions 19 to 21 which concern the reaction:*

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

In the absence of a catalyst, the activation energy for the forward reaction is 183 kJ $mol^{-1}$ and the activation energy for the reverse reaction is 157 kJ $mol^{-1}$.

In the presence of a platinum catalyst the activation energy for the forward reaction is 58 kJ $mol^{-1}$.

**TEE 2000 MC Q19:**

What is the ΔH for the reaction

$$2\,HI(g) \longrightarrow H_2(g) + I_2(g)$$

in the absence of a catalyst?

(a) – 26 kJ $mol^{-1}$
(b) + 26 kJ $mol^{-1}$
(c) – 84 kJ $mol^{-1}$
(d) + 84 kJ $mol^{-1}$

**TEE 2000 MC Q20:**

What is the ΔH for the reaction

$$2\,HI(g) \longrightarrow H_2(g) + I_2(g)$$

in the presence of a platinum catalyst?

(a) – 26 kJ $mol^{-1}$
(b) + 26 kJ $mol^{-1}$
(c) – 84 kJ $mol^{-1}$
(d) + 84 kJ $mol^{-1}$

**TEE 2000 MC Q21:**

Which one of the following statements about the reaction

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

in the presence of a platinum catalyst?

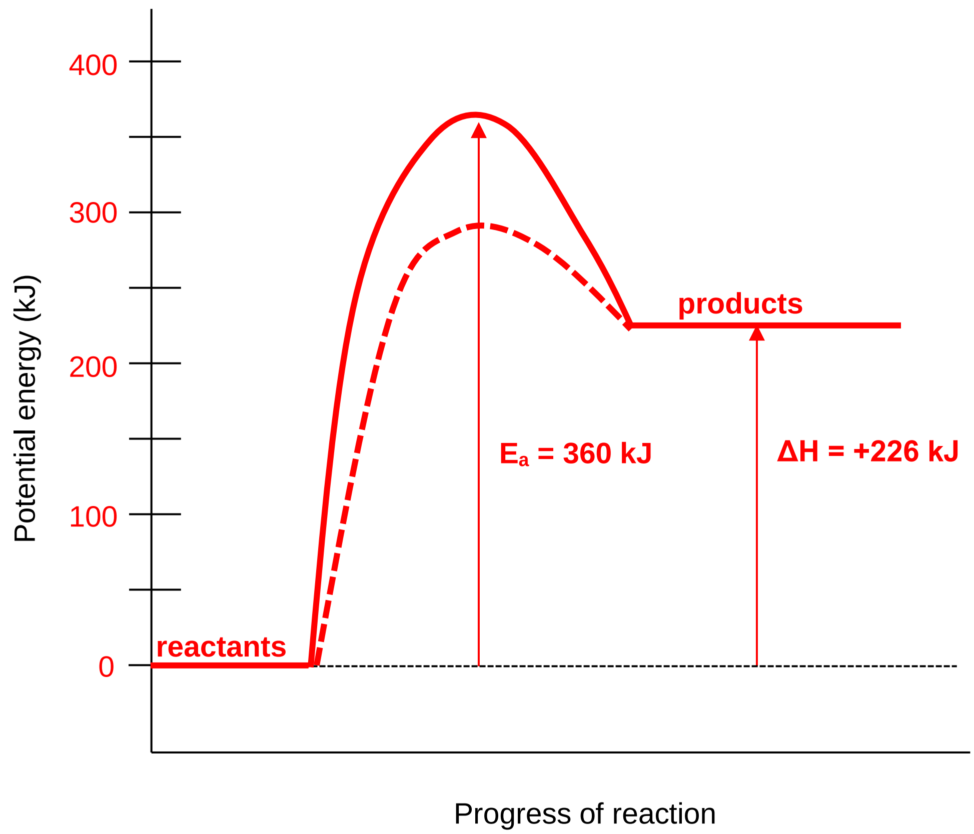
(a) The activation energy is 32 kJ $mol^{-1}$.
(b) The activation energy is 99 kJ $mol^{-1}$.
(c) The activation energy is 157 kJ $mol^{-1}$.
(d) The activation energy cannot be known without further experiments.

**WACE 2012 Q28:**

Consider the following reaction:

$$CO_2(g) + NO(g) \longrightarrow CO(g) + NO_2(g) \qquad \Delta H = +226 \text{ kJ mol}^{-1}, E_a = 360 \text{ kJ}$$

(a) On the axes below draw a potential energy diagram for this reaction. Label the activation energy ($E_a$) and enthalpy change ($\Delta H$) for the reaction. Include a scale on the vertical axis.
On the same axes, use a dashed line to show a possible catalysed pathway. (5 marks)




- **1 mark for appropriate vertical scale**
- **1 mark for shape of graph**
- **1 mark for correctly labelled ΔH**
- **1 mark for correctly labelled $E_a$**
- **1 mark for correct possible catalysed pathway**

(b) i. How much energy is consumed when 2.5 mol of $CO_2(g)$ is reacted with 2.5 mol of $NO(g)$? (1 mark)
**226 x 2.5 = 565 kJ (1 mark)** ***(multiply the energy per mole by the number of moles)***

ii. What is the activation energy when 2.5 mol of $CO_2(g)$ is reacted with 2.5 mol of $NO(g)$? (1 mark)
**360 kJ (1 mark)** ***(this value is the activation energy, regardless of how many moles are reacting)***

**TEE 2008 SA Q7:**

When a piece of freshly cut sodium metal is placed in a jar of chlorine gas (as shown below) it ignites spontaneously with sparks and flashes of flame. A white powder is left on the sides of the jar.

(a) What do the observations tell us about the size of the (2 marks)

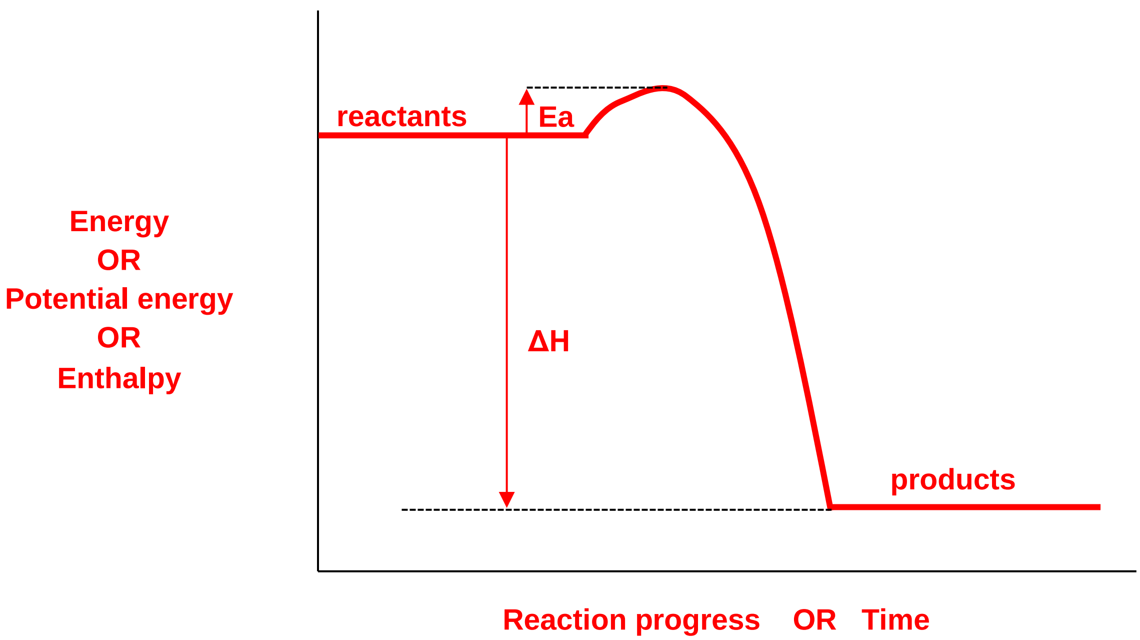
(i) activation energy of this reaction?

**Small (1 mark)** ***(reacts very easily)***

(ii) heat of reaction?

**Large (½ mark), negative (½ mark)** ***(releases a lot of heat energy)***

(b) Draw an energy profile diagram for this reaction, labeling the axes, activation energy and heat of reaction. (3 marks)



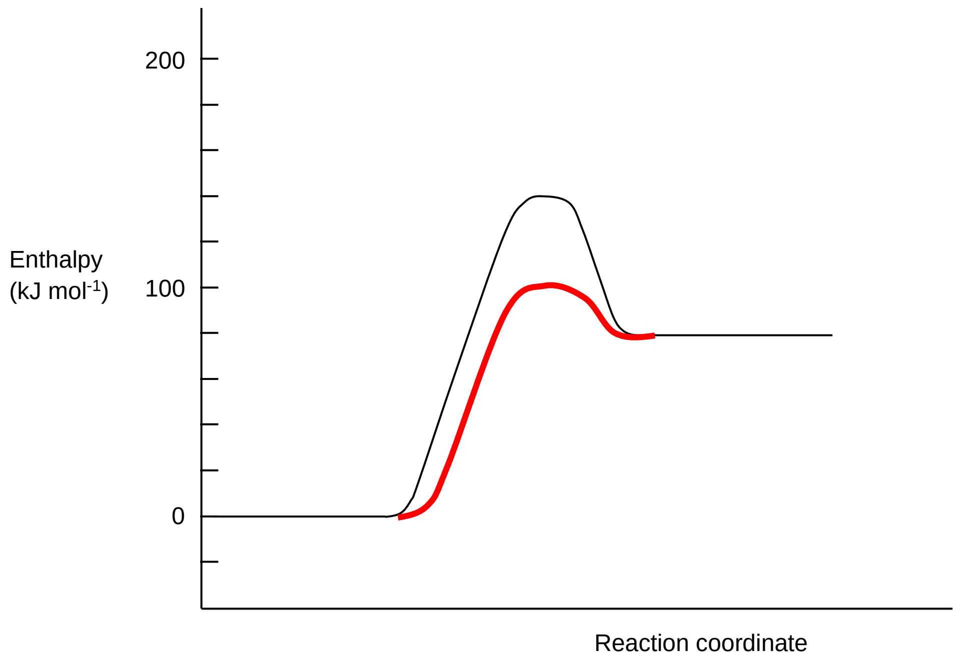

- **1 mark for shape of curve (exothermic, small Ea, large ΔH)**
- **1 mark for axes labels**
- **1 mark for Ea and ΔH labelled correctly**

**TEE 2007 SA Q10:**

Ethene can be produced by heating ethane:

$$C_2H_6(g) \rightleftharpoons C_2H_4(g) + H_2(g)$$

A potential energy diagram for the uncatalysed reaction is shown below:




(a) Determine from this graph the value for the reaction of: (2 marks)

Activation energy: **140 kJ mol$^{-1}$ (1 mark)**

$\Delta H$: **+80 kJ mol$^{-1}$ (1 mark – must <u>explicitly</u> include + sign for 1 mark)**

(b) Using a dotted line, **draw on the above diagram** a possible potential energy diagram for the same reaction but in the presence of a catalyst. (2 marks)

**See graph above**

- **1 mark for lower activation energy**
- **1 mark for same enthalpy of products and reactants**

**TEE 2006 SA Q11:**

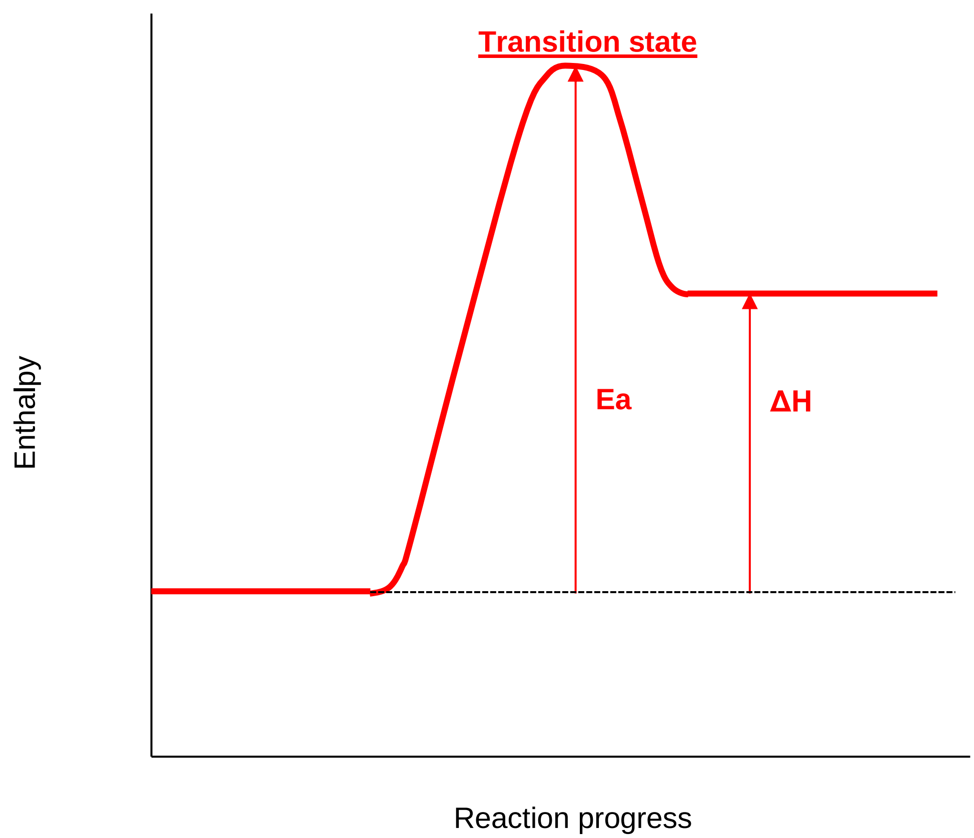
(a) Write the equation for ammonium nitrate dissolving in water. (1 mark)

**$NH_4NO_3(s)$ → $NH_4^+(aq)$ + $NO_3^-(aq)$ (1 mark)**

(b) Given that the reaction is endothermic, describe what you would observe when solid ammonium nitrate is dissolved in a beaker of water (1 mark)

**Beaker becomes cold / Decrease in temperature (1 mark)**

(c) Draw an energy profile diagram to represent this reaction. On your diagram you should include and label the following: activation energy, ΔH, transition state. (4 marks)



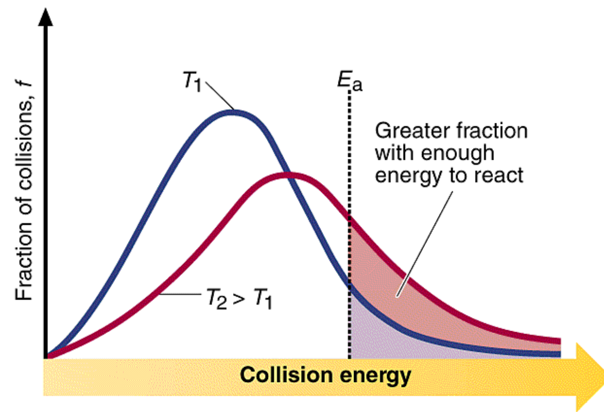

**See graph above**

- **1 mark for shape of graph (endothermic reaction)**
- **1 mark for labelled Ea**
- **1 mark for labelled ΔH**
- **1 mark for transition state** ***(state of reaction with maximum energy)***

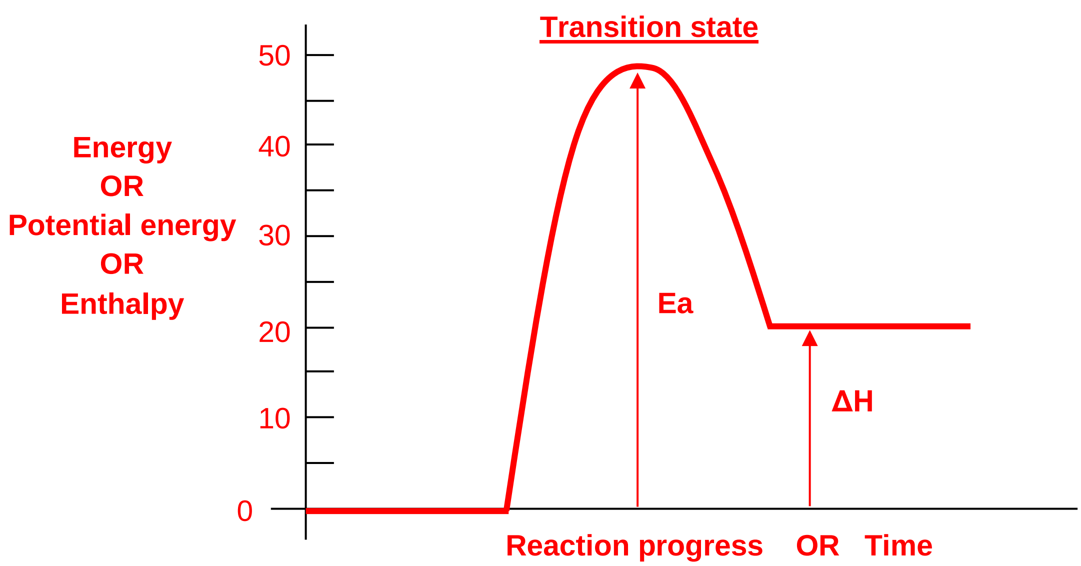
**TEE 2004 SA Q5:**

(a) A small increase in the temperature of a reaction will often cause a significant increase in the rate of a reaction. Explain, with reference to collision theory, why this is so. Use diagrams if appropriate. (4 marks)

- **Particles have greater average kinetic energy (1 mark)**
- **They move faster and collide more often (1 mark)**
- **More particles have energy greater than activation energy, so greater proportion of reactions are successful (1 mark)**
- **Diagram – 1 mark**




(b) Draw a potential energy diagram for a reaction with activation energy = 50 kJ mol$^{-1}$ and ΔH = +20 kJ mol$^{-1}$. Label the diagram **and** axes well, showing the transition state, product and reactants, along with the activation energy and ΔH. (4 mark)




- **Axes labels – 1 mark**
- **Transition state – 1 mark**
- **Activation energy – 1 mark**
- **ΔH and correct graph shape – 1 mark**

**WACE 2016 Sample Q2:** *Originally from WACE 2011*

An enzyme is a biological catalyst. Esters can be hydrolysed, as represented below by an esterase enzyme.

$$\text{ester} + \text{water} \xrightleftharpoons{\text{esterase}} \text{carboxylic acid} + \text{alcohol}$$

In the presence of esterase, which one of the following statements is true for this process?

(a) The position of the equilibrium for this reaction is shifted to the right.
(b) The rate of forward reaction and the rate of reverse reaction both increase equally.
(c) The rate of forward reaction increases more than the rate of reverse reaction.
(d) The rate of forward reaction increases and the rate of reverse reaction decreases.

**WACE 2016 Sample Q3:** *Originally from WACE 2011*

Hydrogen can be produced by the reaction

$$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3\,H_2(g) \qquad \Delta H > 0$$

Which one of the following will increase the equilibrium yield of hydrogen?

(a) increasing the total pressure of the system
(b) decreasing the partial pressure of the water vapour
(c) removing carbon monoxide from the system as it is produced
(d) decreasing the temperature of the system

**WACE 2015 Q2:**

Which one of the following is **true** for a solution of silver chloride in equilibrium with some solid silver chloride, as illustrated by the equation below?

$$Ag^+(aq) + C\ell^-(aq) \rightleftharpoons AgC\ell(s)$$

(a) The silver chloride solution is saturated.
(b) Use of a catalyst would allow more solid silver chloride to dissolve.
(c) If more solid silver chloride is added to the mixture then this will change the concentrations of the silver ions and chloride ions in the solution.
(d) The reaction in which silver ions and chloride ions precipitate to form solid silver chloride is not taking place.

## Section 2: Systems in Equilibrium, Le Châtelier's Principle, Equilibrium expressions

**WACE 2015 Q4:**

Consider the following equilibrium.

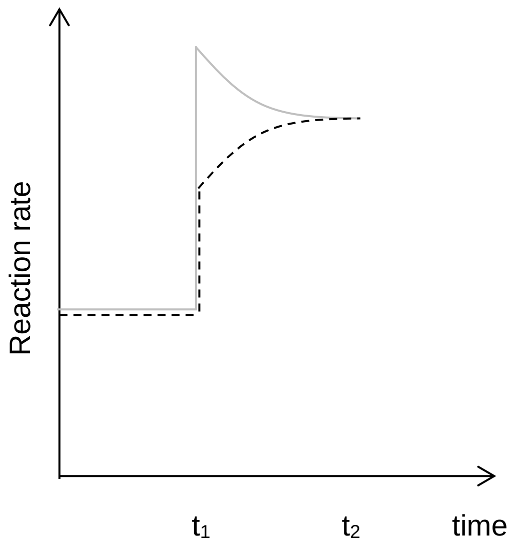
$$2\ C\ell F_3(g) \rightleftharpoons 3\ F_2(g) + C\ell_2(g) \qquad \Delta H = \text{negative}$$

The system is initially at equilibrium. At time $t_1$, the temperature of the system was increased. Which of the following **best** represents the changes in the forward and reverse reaction rates until equilibrium is re-established at time, $t_2$?

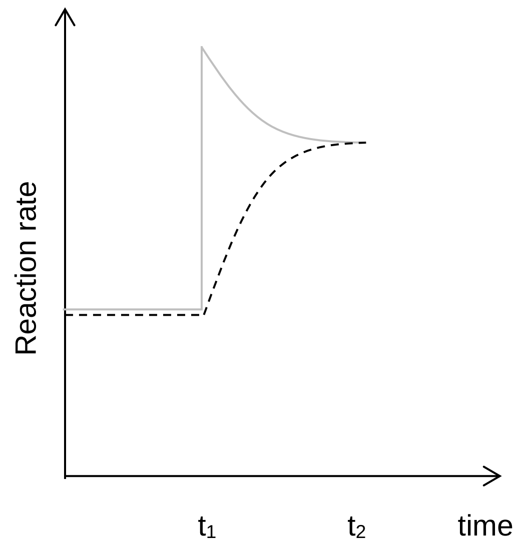
The forward reaction rate is represented by ————————

The reverse reaction rate is represented by - - - - - - - - - - - -

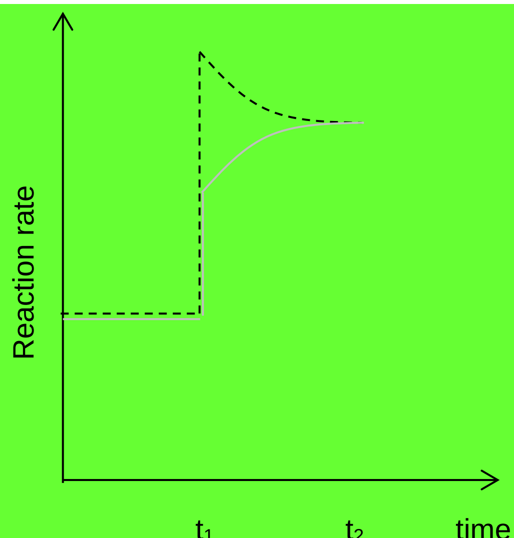
(a)



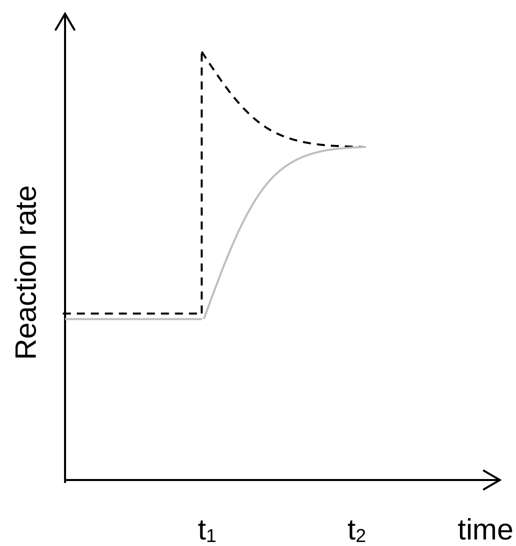

(b)




(c)




(d)



## Section 2: Systems in Equilibrium, Le Châtelier's Principle, Equilibrium expressions

*WACE 2015 Questions 6 and 7 refer to the reaction represented by the equation shown below.*

$$Pb(s) + PbO_2(s) + 4\,H^+(aq) + 2\,SO_4^{2-}(aq) \rightleftharpoons 2\,PbSO_4(s) + 2\,H_2O(\ell)$$

**WACE 2015 Q6:**

Which one of the following is the equilibrium law expression for this reaction?

(a) $K = \frac{1}{[H^+]^2[SO_4^{2-}]^2}$

(b) $K = \frac{1}{[H^+]^4[SO_4^{2-}]^2}$

(c) $K = \frac{[PbSO_4]^2}{[H^+]^4[SO_4^{2-}]^2}$

(d) $K = \frac{1}{[H^+]^2[SO_4^{2-}]}$

**WACE 2015 Q7:**

Assuming equilibrium has been established, which one of the following will cause a decrease in pH?

(a) adding more solid lead
(b) adding solid sodium sulfate
(c) removing solid lead sulfate
(d) adding barium nitrate solution

**WACE 2014 Q10:**

Consider the following endothermic reaction

$$N_2O(g) + NO_2(g) \rightleftharpoons 3\,NO(g) \quad \Delta H = +156 \text{ kJ mol}^{-1}$$

Which one of the following changes to the system will increase the value of its equilibrium constant, K?

(a) increase pressure
(b) addition of a catalyst
(c) increased temperature
(d) decreased temperature

Section 2: Systems in Equilibrium, Le Châtelier's Principle, Equilibrium expressions

**WACE 2014 Q11:**

Which one of the following is the equilibrium law expression for the equilibrium represented below?

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

(a) $\dfrac{[Cr_2O_7^{2-}]}{[CrO_4^{2-}]^2[H^+]^2}$

(b) $\dfrac{[CrO_4^{2-}]^2[H^+]^2}{[Cr_2O_7^{2-}]}$

(c) $\dfrac{[CrO_4^{2-}]^2[H^+]^2}{[Cr_2O_7^{2-}][H_2O]}$

(d) $\dfrac{[Cr_2O_7^{2-}][H_2O]}{[CrO_4^{2-}]^2[H^+]^2}$

**WACE 2014 Q12:**

Aqueous solutions of iron(III) ions and thiocyanate ions form the equilibrium represented below.

| $Fe^{3+}(aq)$ | + | $SCN^-(aq)$ | $\rightleftharpoons$ | $[Fe(SCN)]^{2+}(aq)$ |
|---|---|---|---|---|
| pale brown | | colourless | | deep red |

The reaction is exothermic.

Which one of the following statements about changes to the system and the effect on the colour of the solution is true?

(a) Adding water will make it turn darker red
(b) Cooling the solution will make it turn darker red
(c) Adding a small volume of aqueous $Na_2CO_3$ solution will turn it darker red
(d) Adding solid iron(III) chloride to the solution will make it lighter red

**WACE 2013 Q14:**

Consider the following reaction.

$$2\ SO_2(g) \ + \ O_2(g) \ \rightleftharpoons \ 2\ SO_3(g) \ + \ 198\ kJ$$

After equilibrium has been established, which one of the following would immediately increase the rate of the reverse reaction?

(a) adding a catalyst
(b) increasing the concentration of $SO_2$
(c) cooling the reaction vessel and its contents
(d) adding a small amount of neon gas

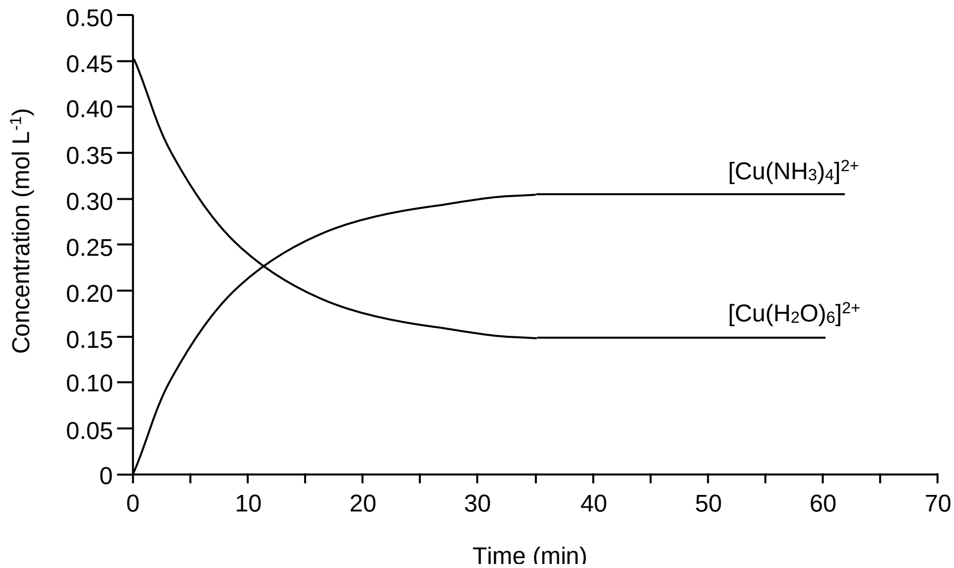
*WACE 2013 Questions 12 and 13 refer to the information and graph below.*

Aqueous solutions of copper(II) ions and ammonia form the equilibrium represented below.

$$[Cu(H_2O)_6]^{2+}(aq) \quad + \quad 4\ NH_3(aq) \quad \rightleftharpoons \quad [Cu(NH_3)_4]^{2+}(aq) \quad + \quad 6\ H_2O(\ell)$$

pale blue deep royal blue

The following graph shows the changes in concentration with time for $[Cu(H_2O)_6]^{2+}$ and $[Cu(NH_3)_4]^{2+}$ ions when solutions of copper(II) nitrate and ammonia are mixed.




**WACE 2013 Q12:**

Which of the following statements is true for this equilibrium system?

(a) The system reaches equilibrium at approximately 35 minutes
(b) At equilibrium, the concentration of $NH_3$ will always be four times greater than the concentration of $[Cu(NH_3)_4]^{2+}$.
(c) Adding ammonia to the system will decrease the equilibrium constant
(d) At equilibrium, the rate of the forward reaction is less than the rate of the reverse reaction.

**WACE 2013 Q13:**

Which one of the following would be observed if a small quantity of concentrated nitric acid was added to the system after it had reached equilibrium?

(a) The solution would turn a deeper royal blue colour.
(b) The solution would be a paler blue colour.
(c) There would be no change in the colour of the system.
(d) Copper(II) nitrate crystals would precipitate from solution.

**WACE 2011 Q12:**

Which of the properties listed below are characteristic of a gaseous system in dynamic equilibrium?

(i) The concentrations of reactants are equal to the concentrations of products
(ii) The concentrations of reactants and products are constant
(iii) The rate of the forward reaction is equal to the rate of the reverse reaction
(iv) The pressure of the system is constant

(a) (i), (ii) and (iii)
(b) (i), (ii) and (iv)
(c) (ii), (iii) and (iv)
(d) (iii) only

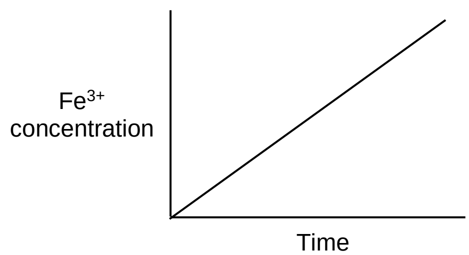
**WACE 2010 Q18:**

When aqueous solutions of $Ag^+$ and $Fe^{2+}$ are mixed, Ag and $Fe^{3+}$ form according to the following equilibrium.

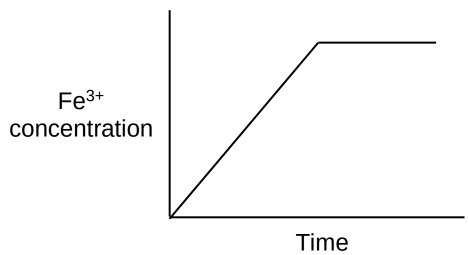
$$Ag^+(aq) + Fe^{2+}(aq) \rightleftharpoons Ag(s) + Fe^{3+}(aq)$$

Which one of the following concentration versus time graphs best represents the way in which the $Fe^{3+}$ concentration varies as the reaction proceeds to equilibrium?

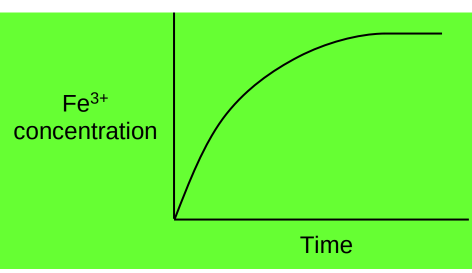
(a)



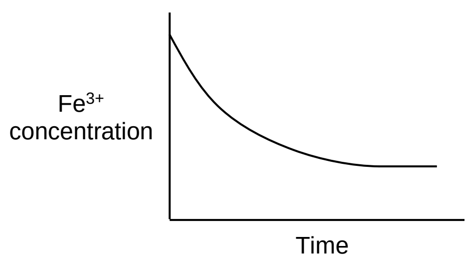

(b)




(c)




(d)



**WACE 2010 Q16:**

Consider the following reaction at equilibrium:

$$2\ SO_2(g) \ + \ O_2(g) \ \rightleftharpoons \ 2\ SO_3(g) \qquad \Delta H = -197\ kJ\ mol^{-1}$$

Which one of the following changes will increase the concentration of $SO_3(g)$ in the mixture when equilibrium is re-established?

(a) decreasing the concentration of $SO_2$ at constant temperature and pressure
(b) decreasing the concentration of $O_2$ at constant temperature and pressure
(c) decreasing the temperature of the system
(d) decreasing the pressure of the system

**TEE 2009 MC Q10:**

The reaction used in the production of ammonia gas is shown below.

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

Addition of a catalyst will increase the rate of this reaction. Which one of the following will occur on the addition of a catalyst?

(a) The equilibrium yield of ammonia remains constant.
(b) The rate of the forward reaction increases relative to the rate of the reverse reaction.
(c) The proportion of successful collisions remains constant.
(d) The endothermic reaction is favoured.

**TEE 2008 MC Q21:**

When solid silver chromate is added to water, the following equilibrium is established:

$$Ag_2CrO_4(s) \ \rightleftharpoons \ 2\ Ag^+(aq) \ + \ CrO_4^{2-}(aq)$$

A small quantity of sodium chromate solid is added to the solution. Assuming there is no change in the volume of the system, which of the following statements is correct?

(a) The concentration of $CrO_4^{2-}(aq)$ will increase and the concentration of $Ag^+(aq)$ will not change.
(b) The concentration of $CrO_4^{2-}(aq)$ will decrease and the concentration of $Ag^+(aq)$ will increase.
(c) The concentration of $CrO_4^{2-}(aq)$ will increase and the concentration of $Ag^+(aq)$ will decrease.
(d) The concentrations of $CrO_4^{2-}(aq)$ and $Ag^+(aq)$ will not change.

**TEE 2008 MC Q22:**

Consider the following system, which is at equilibrium:

$$C_2H_4(g) + HC\ell(aq) \rightleftharpoons CH_3CH_2C\ell(g) \qquad \Delta H = -70 \text{ kJ mol}^{-1}$$

Which of the following statements about this system is true?

(a) The rate of the forward reaction and the rate of the reverse reaction are zero.
(b) The concentrations of the reactants will remain constant over time.
(c) The concentration of $C_2H_4$ will equal the concentration of $CH_3CH_2C\ell$
(d) The sum of the concentrations of $C_2H_4$ and $HC\ell$ will equal the concentration of $CH_3CH_2C\ell$

**TEE 2008 Q23:**

In a chemical reaction at constant temperature, which one of the following statements best describes the result of the addition of a catalyst?

(a) Addition of a catalyst increases the amount of products formed.
(b) Addition of a catalyst decreases the time taken to reach equilibrium.
(c) Addition of a catalyst decreases the amount of energy released in the reaction.
(d) Addition of a catalyst increases the amount of energy released in the reaction.

**TEE 2007 MC Q23:**

A common reaction that illustrates chemical equilibrium is the chromate-dichromate reaction:

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

What is the equilibrium constant for this reaction?

(a) $K = \dfrac{[Cr_2O_7^{2-}][H_2O]}{[CrO_4^{2-}]^2[H^+]^2}$

(b) $K = \dfrac{[Cr_2O_7^{2-}][H_2O]}{[2CrO_4^{2-}][2H^+]}$

(c) $K = \dfrac{[Cr_2O_7^{2-}]}{[CrO_4^{2-}]^2[H^+]^2}$

(d) $K = \dfrac{[Cr_2O_7^{2-}]}{2[CrO_4^{2-}] + 2[H^+]}$

**TEE 2007 MC Q24:**

Which one of the following is characteristic of a system at equilibrium?

(a) The rate of the forward reaction equals the rate of the reverse reaction
(b) The concentration of reactants equals the concentration of products
(c) The forward and reverse reactions have stopped
(d) Changing the temperature of a system in equilibrium has no effect on the equilibrium constant

**TEE 2007 MC Q25:**

If solid calcium carbonate is heated in a sealed container, the following equilibrium is established at 500 °C and 600 kPa pressure:

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g) \qquad \Delta H = +178 \text{ kJ mol}^{-1}$$

Which one of the following statements about this equilibrium is correct?

(a) Adding more $CO_2$ to the system will reduce the amount of CaO present.
(b) Reducing the temperature of the system will increase the amount of CaO present
(c) Increasing the pressure of the system to 1000 kPa by adding inert nitrogen gas will decrease the amount of $CaCO_3$ present.
(d) Adding more $CaCO_3$ to the system will cause in increase in CaO and $CO_2$ present.

**TEE 2007 MC Q26:**

The equilibrium utilised in the Haber process can be represented as:

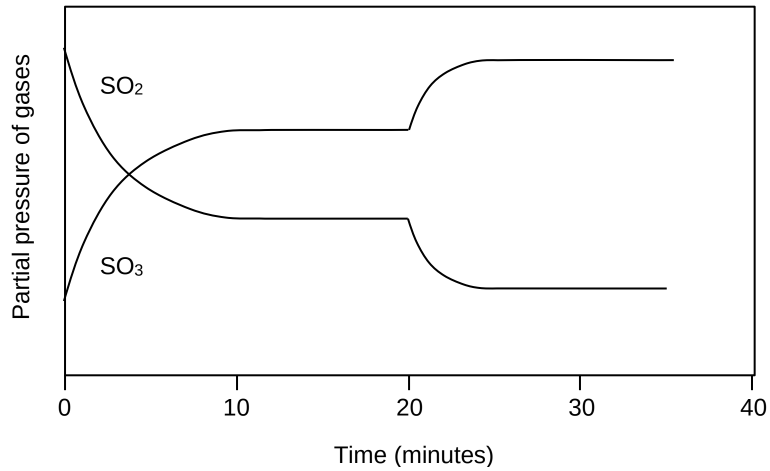
$$N_2(g) + 3\,H_2(g) \xrightleftharpoons{Fe_2O_3 \text{ catalyst}} 2\,NH_3(g) \qquad \Delta H = -92 \text{ kJ mol}^{-1} \quad (\text{at } 25\ °C)$$

What will happen if the quantity of catalyst is halved?

(a) The temperature drops to half the original value.
(b) The rate drops to half the original value.
(c) The yield of product drops to half the original value.
(d) None of the above will occur.

*TEE 2006 Questions 26 and 27 refer to the following graph, which represents the partial pressure of $SO_2$ and $SO_3$ in the reaction shown below.*

$$SO_2(g) + NO_2(g) \rightleftharpoons SO_3(g) + NO(g) \qquad \Delta H = -42 \text{ kJ}$$




**TEE 2006 MC Q26:**

At what time is equilibrium first established?

(a) 5 minutes
(b) 10 minutes
(c) 15 minutes
(d) 30 minutes

**TEE 2006 MC Q27:**

At the 20 minute mark, what changes could have been made to the system to produce the effects shown by the graph?

(a) The system temperature is increased or the partial pressure of NO is increased.
(b) The system temperature is increased or the partial pressure of $NO_2$ is increased.
(c) The system temperature is decreased or the partial pressure of NO is decreased.
(d) The system temperature is decreased or the partial pressure of $NO_2$ is decreased.

## Section 2: Systems in Equilibrium, Le Châtelier's Principle, Equilibrium expressions

*TEE 2006 Questions 28 and 29 refer to the following chemical reaction taking place in a sealed container.*

$$2\ NO(g) + 2\ H_2(g) \rightleftharpoons N_2(g) + 2\ H_2O(g) \qquad \Delta H = -664\ kJ$$

**TEE 2006 MC Q28:**

Which of the following changes made to the system would increase the **equilibrium yield** of $N_2$?

I. Adding a catalyst
II. Increasing the temperature
III. Increasing the pressure
IV. Cooling to cause the $H_2O(g)$ to condense to liquid water

(a) I and II only
(b) III and IV only
(c) II and IV only
(d) I, II and III only

**TEE 2006 MC Q29:**

In the changes referred to in Question 28, which would increase the **rate** of the production of $N_2$?

(a) I and II only
(b) III and IV only
(c) I, II and III only
(d) II, III and IV only

**TEE 2006 MC Q30:**

Consider the reaction:

$$Ca(HCO_3)_2(s) \rightleftharpoons CaO(s) + 2\ CO_2(g) + H_2O(g)$$

Which one of the following is the equilibrium constant expression for this equation?

(a) $K = [CO_2]^2[H_2O]$

(b) $K = 2[CO_2] + [H_2O]$

(c) $K = \dfrac{[CaO][CO_2]^2[H_2O]}{[Ca(HCO_3)_2]}$

(d) $K = \dfrac{1}{[CO_2]^2[H_2O]}$

Section 2: Systems in Equilibrium, Le Châtelier's Principle, Equilibrium expressions

**TEE 2005 MC Q23:**

When silver sulfide is added to water, the following equilibrium is established.

$$Ag_2S(s) \rightleftharpoons 2\,Ag^+(aq) + S^{2-}(aq)$$

The value of the equilibrium constant for this reaction is very small. What does this suggest?

(a) Adding more silver sulfide will increase the amount of ions in solution.
(b) Silver sulfide reacts extensively with water.
(c) The silver sulfide has a very low solubility.
(d) This reaction is endothermic.

**TEE 2005 MC Q24:**

When $CoC\ell_2$ is dissolved in dilute hydrochloric acid, the following equilibrium is established.

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

The solution appears purple in colour as a result of the mixture of the blue and red colours. Which one of the following changes will cause the solution to become more blue in colour?

(a) A catalyst is added.
(b) A few drops of concentrated $HC\ell$ are added.
(c) A few millilitres of $AgNO_3$ solution is added
(d) The solution is diluted by the addition of water.

**TEE 2005 MC Q25:**

If solid calcium carbonate is heated in a sealed container, the following equilibrium is established:

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g) \qquad \Delta H = +178 \text{ kJ mol}^{-1}$$

For this system, which one of the following statements about the equilibrium constant, K, is correct?

(a) K will increase if the pressure of the system is decreased.
(b) K will decrease if the partial pressure of the $CO_2$ is reduced.
(c) K will increase if the temperature of the system is increased.
(d) K will remain constant, regardless of any changes made to the system.

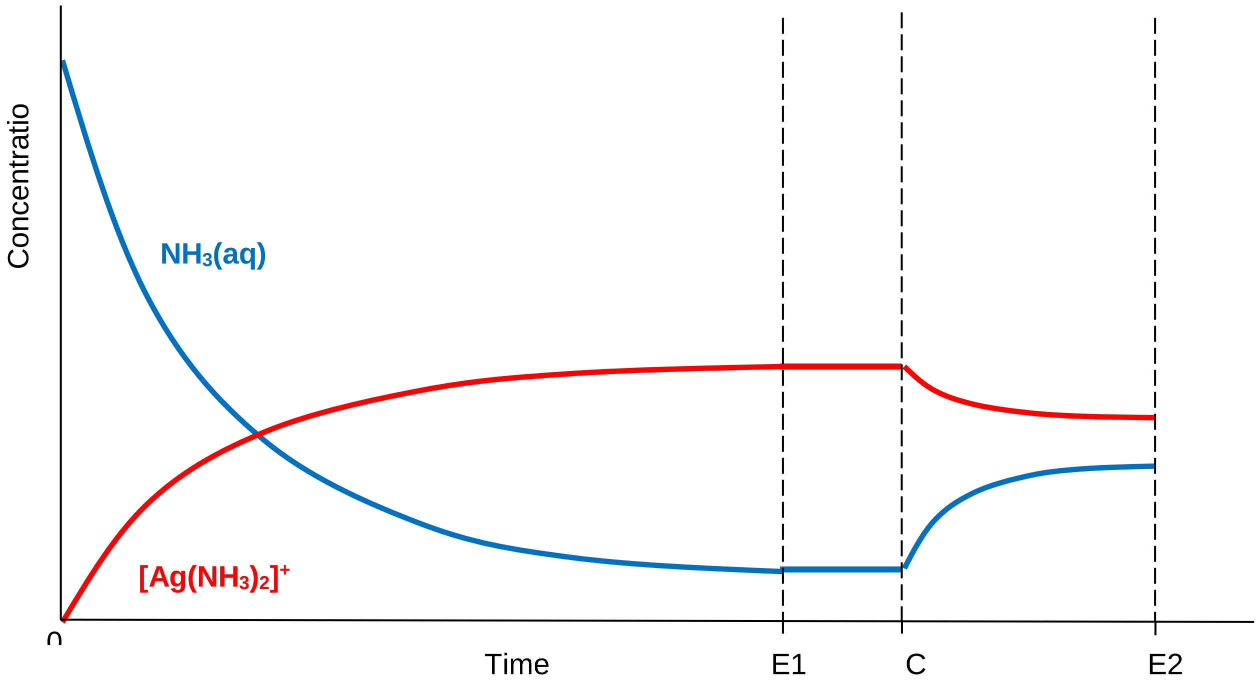
**WACE 2016 Sample Q26:** *Originally from WACE 2012*

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

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

The equilibrium constant, K, for this system is greater than 1 (>1).
A student mixes the reactants at time t = 0.

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




- **Correct shape for both curves, including relative steepness (1 mark)**
- **Straight horizontal lines between E1 and C (1 mark)**
- **Final concentration of $[Ag(NH_3)_2]^+$ > $NH_3$ (1 mark)** ***(needed to match info about K > 1)***

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

- **Correct shape for both curves, including relative steepness (1 mark)**
- **Correct direction of change (1 mark)** ***(adding more $C\ell^-$ forces reaction to the left)***
- **Equilibrium reached <u>at</u> $E_2$ and not <u>before</u> $E_2$. (1 mark)**

**WACE 2016 Sample Q27:** *Originally from WACE 2011*

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

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

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

The value of the equilibrium constant (K) for the above reaction, at 25 °C, is approximately $7.9 \times 10^{-5}$.

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

- **<1 (1 mark) *(low value of K means less products, more reactants)***

(c) Predict the direction in which the equilibrium will shift immediately after the changes indicated in the table below. Write 'left', 'right' or 'no change'. (3 marks)

| Change | Direction of initial equilibrium shift |
|---|---|
| decreasing the temperature | **right (1 mark)** |
| adding hydrochloric acid | **left (1 mark)** |
| adding sodium hydroxide | **right (1 mark)** |

**WACE 2016 Sample Q30:** *Originally from WACE 2011*

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

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

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

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

Write the phase, i.e. solid (s), liquid (ℓ) or gas (g), of each species in this system at the temperatures shown in the table below, and predict the effect of an increase in total pressure on this equilibrium at each of the temperatures. (8 marks)

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

- **4 marks for correctly identifying phases at each temperature**
  ***(1 mark per temperature. All four phases at each temperature must be correct for 1 mark)***

- **4 marks for shifts in equilibrium *(1 mark per temperature)***

  **Note: Shifts in equilibrium must correspond to the phases provided by the student.**

- **WACE 2015 Q30:**

***Note: The marking key for this exam was not released at the time of writing. Answer here is an assumption only.***

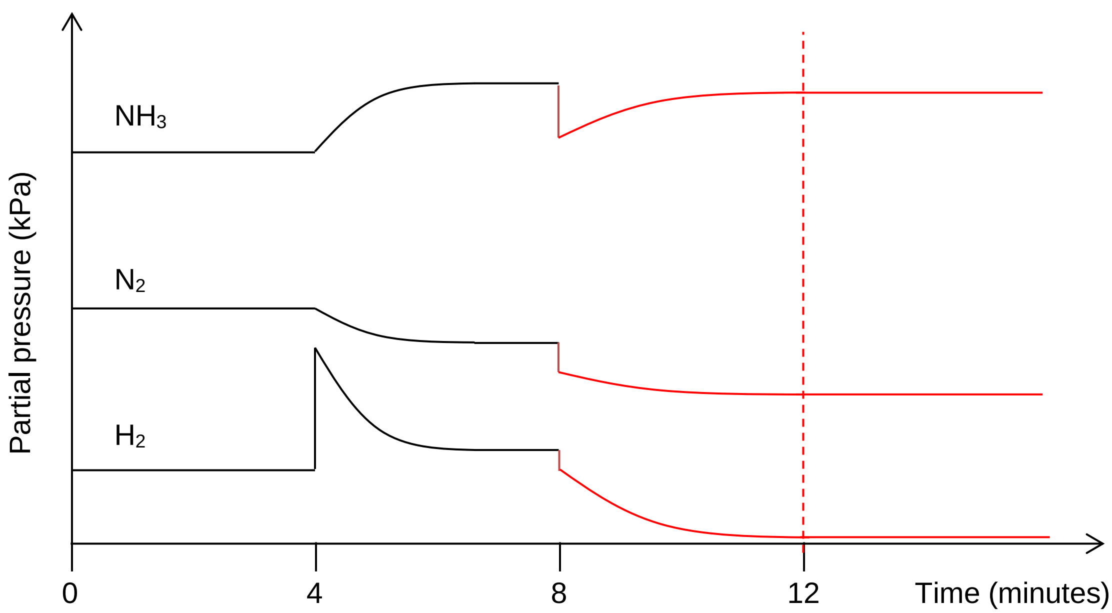
Ammonia exists in equilibrium with hydrogen and nitrogen as shown by the following exothermic equation.

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

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

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

**Partial pressures of $NH_3$, $N_2$ and $H_2$ over time**




(a) What characteristic of equilibrium is indicated on the graph by the section from 0 to 4 minutes?
**constant pressure of gaseous species (1 mark)**

(b) A change was imposed on the system at the 4 minute mark. What imposed change could have produced the results indicated on the graph? (1 mark)
**addition of more $H_2$ *(at constant temperature and volume)* (1 mark)**

(c) The system was **suddenly** cooled at 8 minutes and then reached equilibrium again at 12 minutes. Using this information, complete the graph above from the 8 to the 12 minute mark. (4 marks)
**initial decrease in pressure for all species (1 mark)**
**P($NH_3$) increases over time, P($N_2$) and P($H_2$) decrease over time (1 mark)**
**equilibrium reached at t=12 minutes (1 mark)**
**correct shapes of curves (1 mark)**

**WACE 2013 Q30:**

Consider the following system at equilibrium.

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

Indicate in the table below whether there would be an increase, decrease, or no change in the concentration of $NH_3$(g) after the changes given in the table are imposed on the system and **equilibrium has been re-established**. Provide a brief explanation for the observation. (8 marks)

| Change | Change in concentration of $NH_3$(g) (circle the correct response) | Brief explanation |
|---|---|---|
| The volume of the reaction vessel is doubled | • **decrease** | **Concentrations decrease and rate of forward reaction decreases less than rate of reverse reaction** |
| The temperature of the reaction system is doubled | • **increase** | **Increase in temperature causes the system to move in the direction that consumes heat. Reverse reaction is therefore favoured over forward reaction, increasing the concentration of ammonia.** |
| $N_2$(g) is injected into the reaction system while keeping the volume constant | • **no change** | **($N_2$(g) is not involved in the equilibrium.)**<br><br>**Relative partial pressures of all species remains the same, therefore reaction rates are unchanged.** |
| Water vapour is injected into the reaction system while keeping the volume constant | • **increase** | **Increase in concentration of water vapour increases the rate of reverse reaction relative to the forward reaction.** |

**WACE 2015 Q38:**

***Note: The marking key for this exam was not released at the time of writing. Answer here is an assumption only.***

Two different coloured cobalt(II) complex ions, $Co(H_2O)_6^{2+}$ and $CoCℓ_4^{2-}$, exist together in equilibrium in solution in the presence of chloride ions. This is represented by the equation below.

$$\underset{\text{pink}}{Co(H_2O)_6^{2+}(aq)} + 4\ Cℓ^-(aq) \rightleftharpoons \underset{\text{blue}}{CoCℓ_4^{2-}(aq)} + 6\ H_2O(ℓ)$$

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

| Colour chart | |
|---|---|
| **Species** | **Colour** |
| $Co(H_2O)_6^{2+}(aq)$ | pink |
| $CoCℓ_4^{2-}(aq)$ | blue |
| Initial equilibrium mixture | purple |

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

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

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

| Additions to the test tube | Change in concentration from initial equilibrium to final equilibrium (increase, decrease, unchanged) | | | Colour favoured (pink, blue or unchanged) |
|---|---|---|---|---|
| | $[Co(H_2O)_6^{2+}]$ | $[Cℓ^-]$ | $[CoCℓ_4^{2-}]$ | |
| 1. add $H_2O(ℓ)$ | **Decrease** | **Decrease** | **Decrease** | **Pink** |
| 2. add $HCℓ(aq)$ | **Decrease** | **Increase** | **Increase** | **Blue** |
| 3. add $AgNO_3(aq)$ | **Increase** | **Decrease** | **Decrease** | **Pink** |

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

**White precipitate**

(c) Using Collision theory, explain your predicted observations when hydrochloric acid is added to test tube 2. (3 mark)

**Adding hydrochloric acid increases [$C\ell^-$], leading to more collisions between reactant particles (1 mark)**

**…therefore rate of forwards reaction increases. (1 mark)**

**Forward reaction produces $CoCl_4^{2-}$, so [$CoCl_4^{2-}$] increases and colour changes to blue (1 mark)**

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

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

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

(4 marks)

**Le Chatelier's principle says that a system will shift to partially oppose stress. Therefore cooling a system will favour the exothermic reaction. (1 mark)**

**Pink colour when cooled indicates greater $[Co(H_2O)_6^{2+}]$ (1 mark)**

**Therefore cooling favours the reverse reaction (1 mark)**

**Therefore the reverse reaction is exothermic, and the forwards reaction is endothermic. (1 mark)**

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

**Heavy metals like cobalt can be toxic (1 mark)**
**Waste should be collected for correct disposal, not poured down the sink (1 mark)**

**WACE 2013 Q29:**

Write the equation and the expression for the equilibrium constant for each of the equilibrium processes below. (4 marks)

| Equilibrium process | Equation | Equilibrium constant expression |
|---|---|---|
| Vaporisation of water | **$H_2O(\ell) \rightleftharpoons H_2O(g)$** | $K = [H_2O(g)]$ |
| Dissolution of solid aluminium sulfate in water | **$Al_2(SO_4)_3(s) \rightleftharpoons 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$** | $K = [Al^{3+}]^2 \times [SO_4^{2-}]^3$ |

**WACE 2012 Q42:**

Large public swimming pools are often chlorinated using chlorine gas. The gas is bubbled through the water forming the equilibrium reaction shown below:

$$C\ell_2(aq) + H_2O(\ell) \rightleftharpoons HOC\ell(aq) + H^+(aq) + C\ell^-(aq) \qquad \text{(Reaction 1)}$$

The equilibrium constant for this reaction at 25.0 °C is $3.94 \times 10^4$.

(a) Compare the relative amounts of chlorine and hypochlorous acid ($HOC\ell$) at equilibrium at 25 °C. (1 mark)

**Large value of K indicates more hypochlorous than $C\ell_2$ at equilibrium (1 mark)**

The hypochlorous acid can dissociate as shown in the equilibrium below to give hypochlorite ion.

$$HOC\ell(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + OC\ell^-(aq) \qquad \text{(Reaction 2)}$$

(b) The pH of swimming pools is kept at approximately 7.5. A reason for this is to maximize the concentration of hypochlorous acid, the most effective disinfectant form of chlorine in water. Explain, using the appropriate chemistry concepts, why a pH of about 7.5 will maximize hypochlorous acid concentration. Your answer should consider equilibrium Reactions 1 and 2. (3 marks)

**In Reaction 1, addition of $H^+$ *(acidic conditions)* favours the reverse reaction, reducing $[HOC\ell]$ (1 mark)**

**In Reaction 2, removal of $H^+$ *(basic conditions)* favours the forward reaction, reducing $[HOC\ell]$ (1 mark)**

**An intermediate pH of 7.5 is used because this pH results in the best overall yield of $HOC\ell$ when considering both Reactions 1 and 2 (1 mark)**

**WACE 2012 Q29:**

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

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

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

Complete the table below by indicating the direction of the expected shift in equilibrium immediately following the changes stated in the table. Give the reason for shift. (6 marks)

| Test tube | Change | Direction of shift in equilibrium ('left', 'right' or 'no change') | Reason for shift |
|---|---|---|---|
| A | 3 mL of water is added | **no change** | **Equal number of moles of (aq) ions on each side of the equation** |
| B | A few drops of concentrated nitric acid are added | **right** | **$[H^+]$ increased.**<br><br>**Reaction shifts to decrease the $[H^+]$ so favours forwards reaction** |
| C | A few drops of concentrated silver nitrate solution are added | **right** | **$Ag^+ + C\ell^- \rightarrow AgCl(s)$**<br><br>**Add $Ag^+$ decreases $[C\ell^-]$**<br><br>**Reaction shifts to increase $[C\ell^-]$ which means forwards reaction is favoured.** |

**WACE 2010 Q26:**

Consider the following system:

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

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

| Change | Effect |
|---|---|
| Decreasing the temperature | **decrease** |
| Increasing the pressure of hydrogen | **increase** |
| Adding a catalyst | **increase** |

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

| Change | Effect |
|---|---|
| Increasing the temperature | **decrease** |
| Increasing the pressure of the system | **increase** |
| Adding a catalyst | **no effect** |

**WACE 2010 Q27:**

Write the equilibrium constant expression for the following equilibria:

(a) (1 mark)

| | |
|---|---|
| **Equation** | $BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$ |
| **Equilibrium constant expression** | $K = [Ba^{2+}] \times [SO_4^{2-}]$ |

(b) (1 mark)

| | |
|---|---|
| **Equation** | $2\,CrO_4^{2-}(aq) + 2\,H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + H_2O(\ell)$ |
| **Equilibrium constant expression** | $K = \dfrac{[Cr_2O_7^{2-}]}{[CrO_4^{2-}]^2[H^+]^2}$ |

**TEE 2009 SA Q5:**

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

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

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

**Rates of forwards reaction (dissolving) is equal to rate of reverse reaction (precipitation)**

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

| **Imposed change** | **Effect on solid $Mg(OH)_2$ (write 'increase', 'decrease' or 'no change')** | **Explanation** |
|---|---|---|
| A little concentrated sodium hydroxide solution is added | **increase** | **increasing $[OH^-]$ will increase the rate of the reverse reaction** |
| Some sodium phosphate solution is added to the beaker | **decrease** | **$3\ Mg^{2+} + 2\ PO_4^{3-} \rightarrow Mg_3(PO_4)_2(s)$**<br>**Adding phosphate reduces the $[Mg^{2+}]$.**<br>**This will decrease the rate of reverse reaction, meaning forwards reaction is favoured.** |
| More water is added to the beaker | **decrease`** | **$[Mg^{2+}]$ and $[OH^-]$ will both decrease, leading to a decrease in rate of reverse reaction.**<br>**Forwards reaction will be favoured.** |

**WACE 2010 Q28:**

Ammonia is able to react with itself in the process known as 'self-ionisation'. The equation for the self-ionisation of ammonia is below.

$$NH_3(aq) + NH_3(aq) \rightleftharpoons NH_4^+(aq) + NH_2^-(aq)$$

(b) At standard temperature and pressure, the equilibrium constant, K, for this reaction is about $1 \times 10^{-30}$. The self-ionisation of ammonia is an endothermic process. Will the value of K be less than or greater than $1 \times 10^{-30}$ at temperatures greater than 0 °C? Explain. (3 marks)

**Increasing the temperature of a system will favour the forwards (endothermic) reaction.**

**This will result in an increased [products] and a decreased [reactants]**

**Therefore the value of K will be greater at than $1 \times 10^{-30}$ at higher temperatures.**

**TEE 2007 SA Q6:**

An equilibrium is set up in a test tube by suspending some finely powdered copper sulfide in a dilute solution of hydrochloric acid. The equation for the equilibrium is:

$$CuS(s) + H^+(aq) \rightleftharpoons Cu^{2+}(aq) + HS^-(aq)$$

For each change, list:

- the **immediate** effect on the rate of the forward reaction
- the effect on the yield of $HS^-$ **after equilibrium has been re-established**

Answers should be given as 'increase', 'decrease' or 'no change'. (6 marks)

| Change made to the equilibrium system | Immediate effect on rate of forward reaction | Effect on equilibrium yield of $HS^-(aq)$ |
|---|---|---|
| $HC\ell(g)$ is passed into the solution | **Increase** | **Increase** |
| $CuSO_4$ solution is added | **No effect** | **Decrease** |
| More of the finely powdered CuS is added | **Increase** | **No change** |

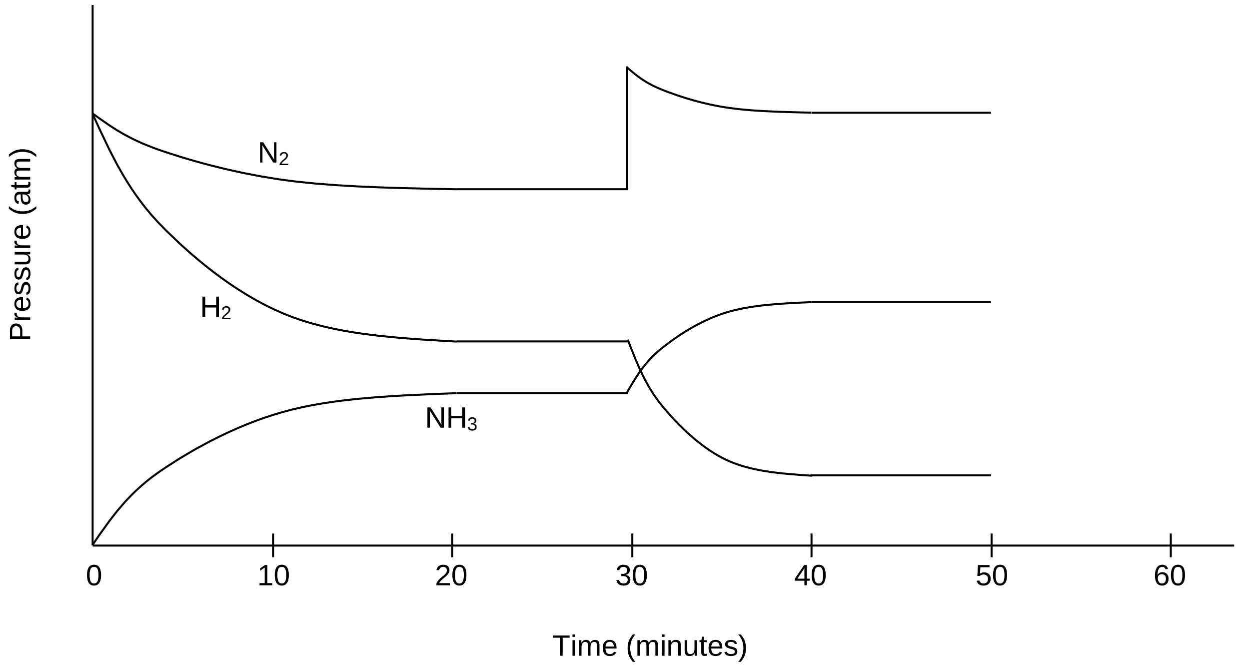
**TEE 2007 SA Q5:**

Ammonia is an industrially important gas produced by the Haber process, as illustrated by the reaction below:

$$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g) \qquad \Delta H = -92\ \text{kJ mol}^{-1}\ \ (\text{at } 25\ °\text{C})$$

The reaction is catalysed by iron(III) oxide, $Fe_2O_3$.

The following graph shows the partial pressures of the three species involved in the reaction.




Answer the following questions about the above graph.

(a) Why does the partial pressure of the $H_2$ decrease more rapidly than that of the $N_2$? (1 mark)

**3 moles of $H_2$ are consumed for each mole of $N_2$ consumed**

(b) Why do the partial pressures of each of the three species stabilise between 20 and 30 minutes? (1 mark)

**System has reacted equilibrium**

(c) What has occurred at the 30-minute mark to cause the changes shown in the graph? (1 mark)

**More $N_2$ is added to the system**

(d) By the 40-minute mark, what difference will the change imposed at the 30-minute mark have made to the rate of: (2 marks)

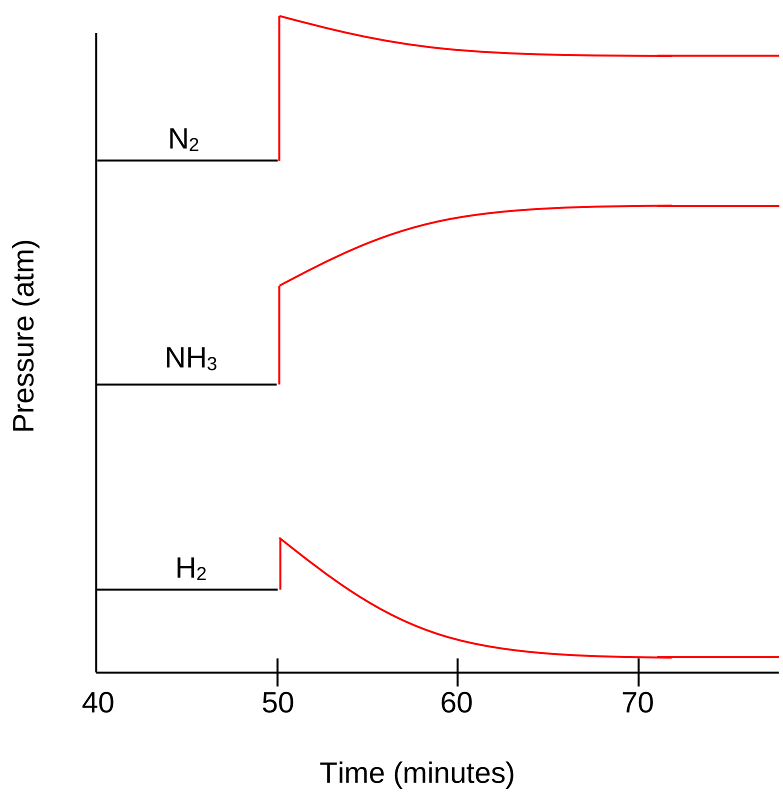
the forward reaction? **increased**

the reverse reaction? **increased**

(e) Using the Collision Theory, explain why the rate of forward reaction is affected by the imposed change at the 30-minute mark. (2 marks)

**Greater [$N_2$] leads to more collisions between $N_2$ and $H_2$ molecules, therefore greater reaction rate**

(f) At 50 minutes, the contents of the reaction vessel are rapidly compressed by reducing the volume. The changes in the partial pressures of the species are shown on the following graph, starting at 40 minutes.




Complete the above graph up to 70 minutes by shown how the partial pressures of each of the species change as a new equilibrium is achieved. (3 marks)

- **Initial increase in pressure (1 mark)**
- **Increasing [$NH_3$] and decreasing [$N_2$] and [$H_2$] in correct ratios (1 mark)**
- **Equilibrium @ t=70 minutes. (1 mark)**

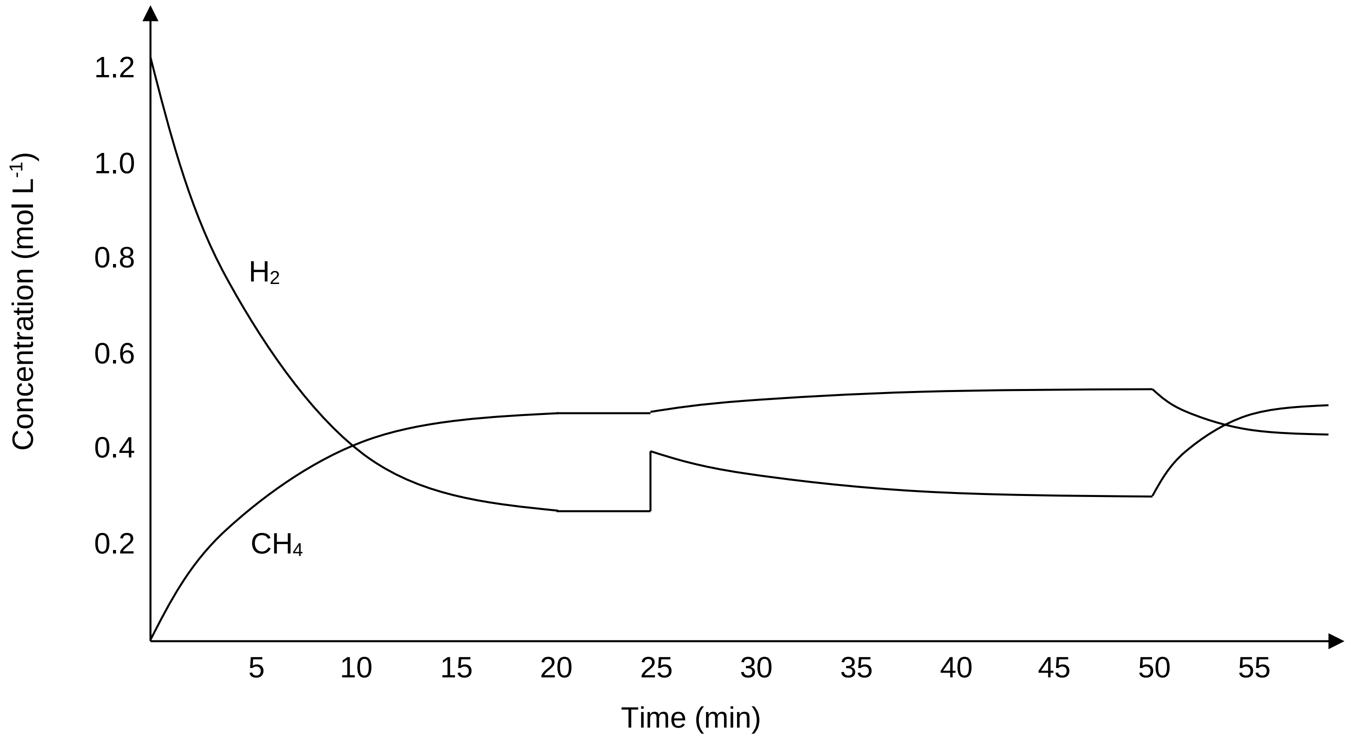
**TEE 2005 SA Q7:**

The reaction between carbon and hydrogen gas to form methane can be represented by the following equation.

$$C(s) \ + \ 2\,H_2(g) \ \rightleftharpoons \ CH_4(g) \ + \ 75\text{ kJ}$$

***exo***
$\rightleftharpoons$
***endo***

The concentrations of hydrogen and methane were plotted over time and the following graph produced.




(a) What time was equilibrium first established? (1 mark)

**20 minutes**

(b) Suggest what could have caused the change at the 25 minute mark. (1 mark)

**Addition of $H_2$ to the vessel**

(c) Suggest what change to the system occurred at the 50 minute mark. (1 mark)

**Increase in temperature** ***(favours the endothermic reverse reaction)***

(d) What would be the effect on the equilibrium if more C(s) was added to the system? (1 mark)

**No effect**

(e) Predict, using Le Châtelier's Principle, what would be the effect of having the volume of the reaction container. (2 marks)

**Halving the volume will increase overall pressure.**

**System will shift in a way that will reduce overall pressure (1 mark)**

**Will favour the formation of products because this side of the equation has less moles of gas (2 moles → 1 mole) (1 mark)**

**TEE 2006 SA Q9:**

When chlorine gas is added to water, the following equilibrium is established:

$$C\ell_2(g) + H_2O(\ell) \rightleftharpoons HOC\ell(aq) + H^+(aq) + C\ell^-(aq) \qquad \Delta H = +ve$$

*endo* $\rightleftharpoons$ *exo*

(a) Write the equilibrium constant expression for this reaction (2 marks)

$$K = \frac{[HOCl][H^+][Cl^-]}{[Cl_2]}$$

(b) Complete the following table. Answers should be given as "increases", "decreases" or "no change". (8 marks)

| Change made to the equilibrium system | Immediate effect on rate of forward reaction | Effect on equilibrium yield of HOCℓ(aq) |
|---|---|---|
| Increase the partial pressure of $C\ell_2(g)$ | **Increase** | **Increase** |
| Increase the temperature of the system | **Increase** | **Increase** |
| Acidify the solution by the addition of nitric acid solution | **No effect** *(on initial rate of forwards reaction)* | **Decrease** |
| Add a suitable catalyst | **Increase** | **No effect** |

**TEE 2008 SA Q6:**

The following equilibrium is set up by adding solid silver chloride to dilute ammonia solution in three test tubes:

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

(a) Write an equilibrium constant expression for this equation. (1 mark)

$$K = \frac{[Ag(NH_3)_2^+][C\ell^-]}{[NH_3]^2}$$

(b) The following changes are made to the equilibrium system. Each change is applied to a separate test tube and equilibrium is re-established. Complete the table below, indicating the changes in the forward reaction rate, and the concentration of $Ag(NH_3)_2^+(aq)$ compared to the original equilibrium system. Use the terms 'increase', 'decrease' or 'no change'.

Also describe what you would observe as equilibrium is re-established in the system.

| | At new equilibrium | | |
|---|---|---|---|
| **Imposed change** | **Effect on reaction rate** | **Effect $[Ag(NH_3)_2^+](aq)$** | **Observation** |
| $NH_3(g)$ is bubbled through the solution | **Increase** | **Increase** | **Some solid dissolves** |
| $NaC\ell(s)$ is added to the solution | **Increase** | **Decrease** | **More solid forms** |
| A few drops of concentrated $HNO_3(aq)$ are added to the solution. | **Decrease** | **Decrease** | **More solid forms** |

**WACE 2014 Q30:**

Hydrogen can be made by reacting methane (natural gas) with water (steam). The reaction can form the chemical equilibrium represented below.

$$CH_4(g) + H_2O(g) \rightleftharpoons 3\ H_2(g) + CO(g) \qquad \Delta H = +206\ kJ\ mol^{-1}$$

State the conditions of temperature and pressure that would optimize the yield of hydrogen at a reasonable rate of reaction. Using collision theory and principles of chemical equilibrium, explain your choice of conditions.

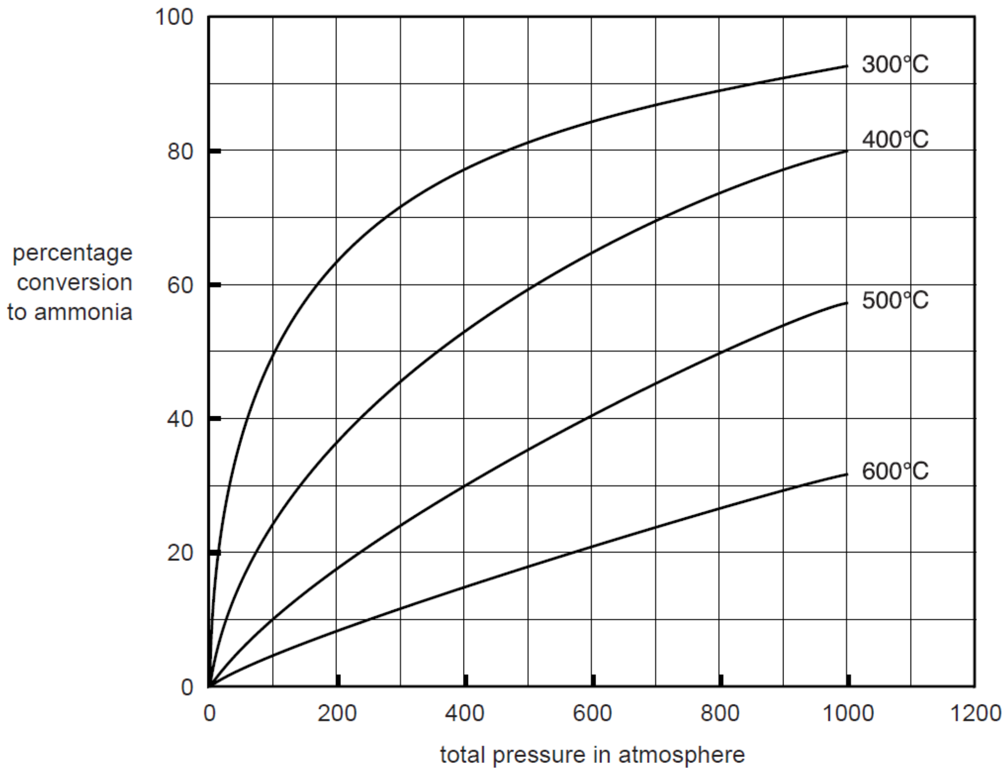
| | **Optimum conditions** | **Explanation** |
|---|---|---|
| Temperature | **high**<br><br>(1 mark) | **Example answer for full marks:**<br>*High temperature increases the proportion of molecules colliding with energy above the $E_a$ and so increasing the reaction rates for both the forward and reverse reactions but the (forward) endothermic direction will increase more so increasing yield of $H_2$.*<br><br>**Mark breakdown:**<br>• Recognition that high temperature increases the proportion of molecules colliding with energy above the $E_a$ (1)<br>• Recognition that high temperature increases rates of forward (and reverse) reaction(s) (1)<br>• Recognition that high temperature increases the rate of forward reaction more than rate of reverse reaction **or** accept Le Chatelier's Principle explanation (1) |
| Pressure | **moderate**<br><br>(1 mark) | **Example answer for full marks:**<br>High pressure increases frequency of collisions between molecules and increases rates for both the forward and reverse reactions but increases reverse reaction rate more (because there are fewer gas molecules on reactant side). Low pressure will increase yield of $H_2$ but the rate of the reaction will be too slow so a compromise moderate pressure is needed.<br><br>**Mark breakdown:**<br>• Recognition that high pressure increases frequency of collisions between molecules (1)<br>• Recognition that high pressure increases rates of **both** forward and reverse reactions (1)<br>• Recognition that high pressure increases rate of reverse reaction more than forward (1)<br>• Recognition that low pressure will increase yield of $H_2$ but the rate of the reaction will be too slow. Compromise between yield and reaction rate (1) |

**VCE 2002 Question 4:**

Ammonia is prepared industrially from hydrogen and nitrogen in the presence of a suitable catalyst according to the equation:

$$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$$

The graph below shows the variation of the equilibrium yield of ammonia with pressures at different temperatures.




(a) A particular industrial plant uses a pressure of 300 atm and a temperature of 500 °C. From the graph, determine the percentage yield of ammonia under these conditions. (1 mark)

**24% (± 1%) (1 mark)**

(b) State Le Châtelier's principle. (2 marks)

**A system at equilibrium will shift in a way to partially counteract the effect of an imposed change.**

Deduce from the graph whether the production of ammonia from hydrogen and nitrogen is an exothermic or an endothermic reaction. Explain your reasoning. (2 marks)

**Forwards reaction is favoured at low temperatures, as seen by higher yield of $NH_3$ (1 mark)**

**Low temperatures favour the exothermic reaction, therefore forwards reaction is exothermic (1 mark)**

(c) Temperatures less the 400 °C are not used for this industrial reaction even though such temperatures give a greater equilibrium yield of ammonia. Give a possible reason why this is so. (1 mark)

**Rate may be too slow at temperatures < 400 °C.**

**WACE 2016 Sample Q34:**

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

(a) Write a balanced equation for this equilibrium. (2 marks)

$CO_2(g) + H_2O(\ell) \rightleftharpoons H_2CO_3(aq)$

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

$H_2CO_3(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$

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

**Acidification affects the process of calcification for marine organisms.**

**At lower pH conditions, calcium carbonate is more likely to dissolve**

$H_3O^+(aq) + CaCO_3(s) \rightarrow Ca^{2+}(aq) + HCO_3^-(aq) + H_2O(\ell)$

**WACE 2016 Sample Q42:**

Use the following information about bleaching to answer the questions that follow.

Many chlorine-based compounds, such as sodium hypochlorite ($NaOC\ell$), chlorine ($C\ell_2$) and chlorine dioxide ($C\ell O_2$), are used as bleaches in household cleaning products and for industrial processes.

Their uses include:

- removing colour (for example, stain removal from clothes)
- whitening paper pulp in the process of making paper
- sterilizing substances (for example, swimming pool water).

These compounds act by oxidising the compounds with which they come into contact. When chlorine gas is used for bleaching, the active ingredient is hypochlorous acid ($HOC\ell$). This is produced by reaction of the chlorine gas with water. Hydrochloric acid is also produced in the reaction.

To increase the amount of hypochlorous acid produced in this reaction, the water through which the chlorine is bubbled is usually made alkaline by the addition of a small amount of hydroxide ions. chlorine-based bleaches react well at room temperatures.

A disadvantage of chlorine bleaches is the potential for highly poisonous dioxins to be produced by reaction with organic compounds. Peroxide bleaches are environmentally more acceptable because they produce oxygen and water.

Hydrogen peroxide is a liquid, but sodium percarbonate ($2Na_2CO_3{\cdot}3H_2O_2$) and sodium perborate ($NaBO_3{\cdot}H_2O$) are solid peroxide bleaches that release hydrogen peroxide when dissolved in water. A disadvantage of peroxide bleaches is the need for high temperatures for them to react.

The development of molecules known as tetra-amido macrocyclic ligand-activators (TAMLs) that function as catalysts has enabled the hydrogen peroxide bleaching reaction to occur at much lower temperatures.

(a) Write the balanced equation for the reaction of chlorine gas with water. (1 mark)

**$C\ell_2(g) + H_2O(\ell) \rightleftharpoons HC\ell O(aq) + HC\ell(aq)$**

(b) Explain briefly how the addition of hydroxide ions to the water through which the chlorine is bubbled will increase the amount of hypochlorous acid produced. (3 marks)

- **Hydroxide ions react with $HC\ell$ and reduce [$HC\ell$] (1 mark)**
- **This slows down the reverse reaction… (1 mark)**
- **…making the forwards reaction favoured and leading to greater yield of $HOC\ell$ (1 mark)**

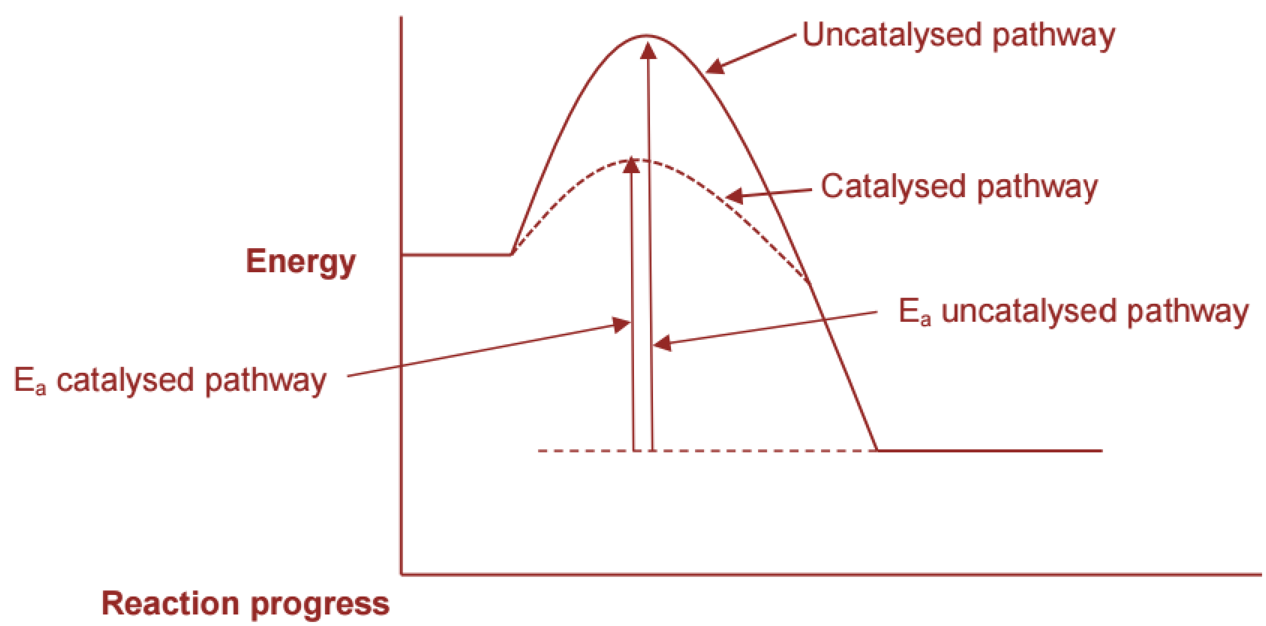
(g) Compare the activation energy for oxidation reactions involving chlorine-based bleaches to those using peroxide-based bleaches (in the absence of catalysts). Explain the reasons for your answer.
(2 marks)

**Chlorine-based bleaches have less Ea than peroxide-based bleaches**

**Peroxide-based bleaches need higher temperatures to react, indicating that there is a larger activation energy barrier that needs to be overcome.**

(h) Draw a fully labeled energy profile diagram showing the progress of the decomposition of hydrogen peroxide with and without TAML molecules.

The equation for the reaction is $2\ H_2O_2 \rightarrow 2\ H_2O + O_2 +$ energy. (4 marks)



**VCE 2004 Q5:**

The industrial production of sulfuric acid can be described as a four-stage process beginning with the burning of raw sulfur with oxygen.

(a) **Stage 1: The burning of sulfur**
Give the equation for the burning of sulfur in oxygen. (1 mark)

**$S(s) + O_2(g) \rightarrow SO_2(g)$**

(b) **Stage 2: The oxidation of sulfur from $SO_2$ to $SO_3$.**
i. Give the equation for this reaction. (1 mark)

**$SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g)$** ***Note: Reversible arrow should be used in equation***

ii. What goes wrong in the industrial process if the temperature for this stage of the process is too high, and why? (2 marks)

**High temperatures would result in less equilibrium yield of $SO_3$.**

**The forwards reaction is exothermic. According to Le Chatelier's Principle, increasing the temperature will favour the endothermic reaction (i.e. the reverse reaction).**

iii. What goes wrong in the industrial process if the temperature for this stage of the process is too low, and why? (2 marks)

**Low temperatures would result in a slow rate of reaction.**

**Decreasing the temperature means that particles have less kinetic energy. They will collide less often and less particles will more energy than the activation energy, so less successful collisions.**

(c) **Stage 3: The conversion of $SO_3$ from a gaseous form to a liquid form by reacting the gas with a suitable solvent.**

i. Give the chemical equation for this process. (1 mark)

**$SO_3(g) + H_2SO_4(\ell) \rightarrow H_2S_2O_7(\ell)$**

ii. Explain why water is not used as a solvent for this process. (1 mark)

**The reaction would be uncontrollable and would create a fog of sulfuric acid.**

(d) **Stage 4: The production of liquid sulfuric acid**

Give the chemical equation for this process. (1 mark)

**$H_2S_2O_7(\ell) + H_2O(\ell) \rightarrow 2\ H_2SO_4(\ell)$**

**WACE 2010 Q42:**

Nitric acid is manufactured by the Ostwald process.

In the first step, ammonia gas reacts with oxygen gas to produce nitric oxide in the presence of a catalyst such as platinum with 10% rhodium. This reaction is carried out at a temperature of approximately 900 °C and at a pressure of approximately 10 atmospheres.

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

The nitric oxide is next oxidised at approximately 50 °C.

$$2\ NO(g)\ +\ O_2(g)\ \rightleftharpoons\ 2\ NO_2(g)\ +\ \text{heat}$$

The nitrogen dioxide then enters an absorption tower, where water is added through a sprinkler system in the presence of air to give nitric acid.

$$4\ NO_2(g)\ +\ O_2(g)\ +\ 2\ H_2O(\ell)\ \rightleftharpoons\ 4\ HNO_3(aq)$$

Use your understanding of reaction rates and chemical equilibrium to explain the conditions used in the Ostwald process. Your answer should include at least three (3) paragraphs, and should be 1 to 1½ pages in length. (10 marks)

## SAMPLE ANSWER:

*The following is an actual answer from the 2010 WACE Examination. This answer received a mark of 10/10 by the WACE examiner's that year. This is the level that is expected under exam conditions.*

"For the first step, a catalyst is used. This will increase the reaction rate by producing an alternative pathway where the activation energy is lower and therefore more collisions will have $E_k > E_a$ and reaction rate is increased. Because it increases both the forward and reverse reaction it won't affect the equilibrium yield.

High temperature is used. This will increase the $E_k$ of the particles, therefore more collisions have $E_k > E_a$ and reaction rate is increased. But it will decrease the equilibrium yield. As the reverse reaction consumes heat, the system will be shifted to the left when temperature increases. Therefore a compromising temperature will need to be found, and in this case, 900 °C, which is a moderate high temperature.

High pressure is also used in this step. This will increase the concentration of the particles and more collisions will occur. ∴ Increase the reaction rate. But this will decrease the equilibrium yield, as the reactant side has less gas molecules (9) than product side (10), and when pressure increases the system will be shifted to the side with less molecules. Therefore a compromising pressure will need to be found. In this case, 10 atm, which is a moderate high pressure.

The temperature and pressure are actually high. This might be because the reaction has very high equilibrium yield as S.T.P., and chemical engineers are concentrating on increasing the reaction rate.

Step 2 used low temperature. This will increase the equilibrium yield because the forward reaction produces heat. But will decrease the reaction rate. A possible explanation is that maybe the yield is too low, and chemical engineers are trying to increase it.

Step 3, water is added through a sprinkler system. Because this is a reaction between gas and liquid, collisions only occur at the surface. ∴ using a sprinkler system could increase the surface area of water, and more collisions will occur and increase the reaction rate."

## MARKING GUIDE:

**The candidate must give an expansive answer. For example, a statement that a reaction rate increases with increasing temperature should be supported by an explanation of why this is the case.**

**A student must address:**

- **temperature, pressure and catalyst in step 1 (6 marks)**
- **temperature in step 2 (2 marks)**
- **water drops in step 3 (2 marks)**

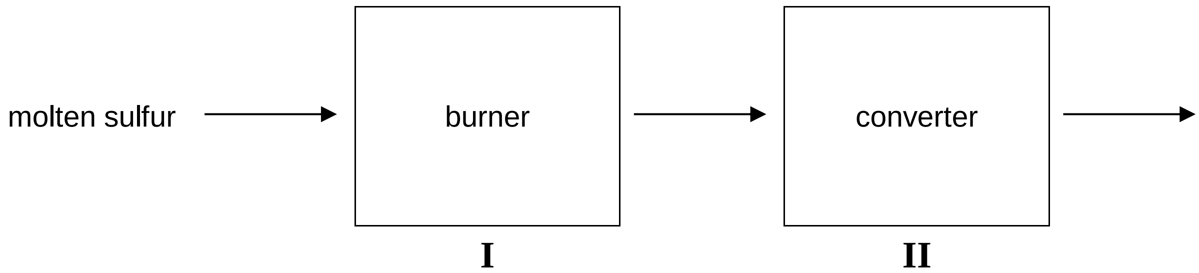
**Each condition must be addressed from a rate and equilibrium perspective**

## COMMENTS FROM EXAMINERS:

This question was not well done, and highlighted the difficulties students often have in constructing coherent sentences and clearly conveying their understanding of a concept in writing. Many candidates discussed things tangential to the question and discussed and introduced things not given, or not relevant to, the question. Some students restated the information given in the question; this is not a productive use of time – no marks are gained for simply restating information provided in the question. This practice should be discouraged. The examining panel urges teachers to encourage candidates to be succinct and direct in their answers to these types of explanatory questions. There is evidence that a concise answer, in general, receives a better score than an answer that is not concise.

**VCE 2005 Question 5:**

Sulfuric acid can be produced from mined sulfur via the Contact Process. The first two stages in the industrial production of sulfuric acid by this process are represented below.




(a) Give a reason why, in stage I, the molten sulfur is sprayed into the burner rather than being allow to flow through it. (1 mark)

**Spraying the molten sulfur produces droplets which provide a greater surface area for contact with $O_2$ and hence speed up the reaction.**

(b) A conflict is involved in choosing the best temperature to be used in stage II, where the reaction is:

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

i. Describe the nature of the conflict and explain how the conflict is resolved. (2 marks)

**Because the forward reaction is exothermic, the yield of $SO_3$ is greater at lower temperatures. However, the reaction is faster at higher temperatures. (1 mark)**

**The rate/yield conflict is resolved by using a temperature which provides a balance (compromise) between the conflicting effects. (1 mark)**

ii. Would increasing the pressure of the reacting mixture in the converter affect the amount of $SO_3$ produced in stage II? Explain your answer. (2 marks)

**The amount of $SO_3$ will increase (1 mark)**

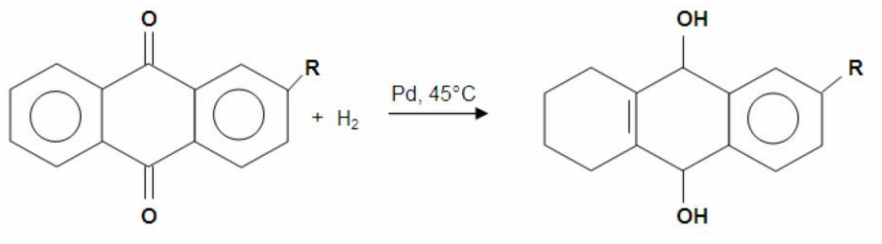
**The equilibrium responds to pressure increases by moving to the side with fewer moles of particles. There are less moles of gas on the product side (2 moles) than the reactant side (3 moles). (1 mark)**

**WACE 2010 Sample Q39:**

Hydrogen peroxide ($H_2O_2$) is an important industrial oxidizing agent. Its manufacture includes a number of steps, the first of which involves the hydrogenation of an alkyl anthraquinone.

**Step 1 – Hydrogenation**

Hydrogen gas is bubbled through a solution containing an alkyl anthraquinone in two solvents, one polar (in which very little anthraquinone dissolves) and the other non-polar. Finely divided alumina particles loaded with palladium catalyst are added to the solution. A number of hydrogenation reactions occur to convert the alkyl anthraquinone (**1**) into tetrahydro-alkyl anthrahydroquinone (**2**) as shown below. The palladium catalyst is removed by filtration before step 2.




1 2

(b) In the hydrogenation step of this process, what effect does the palladium have on the rate at which equilibrium is attained? Explain, by applying Collision Theory, how the palladium has this effect. (3 marks)

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(c) Explain why the palladium in the hydrogenation step is finely divided. (2 marks)

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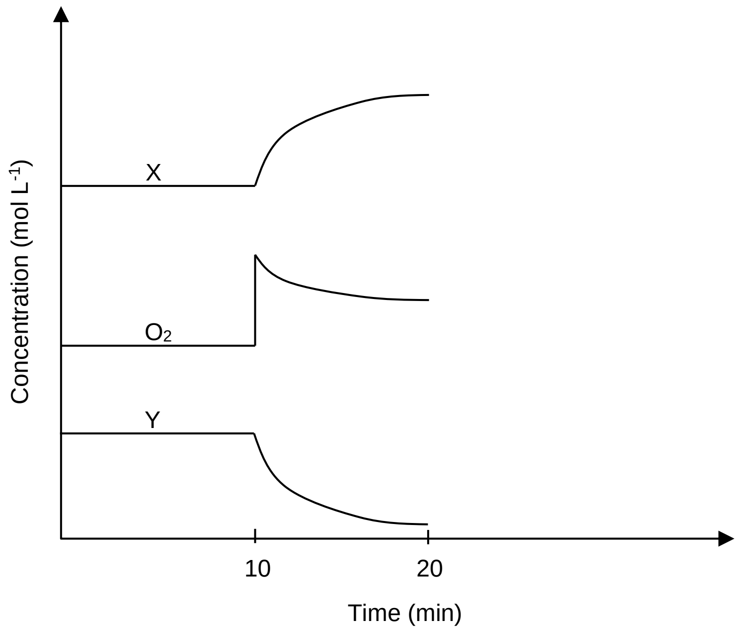
**VCE 2003 Q2:**

Part of the Contact Process for the manufacture of sulfuric acid involves the conversion of sulfur dioxide to sulfur trioxide, as shown by the equation

$$2\ SO_2(g) \ + \ O_2(g) \ \rightleftharpoons \ 2\ SO_3(g) \qquad \Delta H = -192\ kJ\ mol^{-1}$$

As part of a laboratory study for this process, a container was filled with an equilibrium mixture of sulfur dioxide, sulfur trioxide and oxygen in the presence of a catalyst. The container was initially at 450 °C. The container had a fixed volume and was **thermally well insulated**.

Concentrations during a following experiment are shown on the diagram below.




(a) Give the name or formula of a possible catalyst that was added to the reaction mixture. (1 mark)

**vanadium(V) oxide OR vanadium pentoxide**

(b) What change occurred at the 10 minute point? (1 mark)

**Some $O_2$ was added to the container**

(c) Which components of the equilibrium mixture are represented by X and Y? (1 mark)

X = **$SO_3$**

Y = **$SO_2$**

(d) Give explanations for the changes in concentration that occur in X, Y and $O_2$ between 10 and 20 minutes. (3 marks)

**Increasing the concentration of $O_2$ causes the rate of forward reaction to increase due to more collisions between reactant particles.**

**As forward rate > reverse rate:**

- **the concentration of reactants ($O_2$ and X/$SO_3$) decrease over time**
- **the concentration of products (Y/$SO_2$) increase over time**

**This continues until the increasing [products] and decrease [reactants] cause the forward and reverse rates to become equal again at t=20 minutes.**

(e) Would the temperature of the mixture **increase, decrease** or **remain the same** between 10 and 20 minutes? Explain your reasoning. (2 marks)

**The temperature will increase (1 mark) because the forward reaction is exothermic, and this was the favoured reaction between 10 and 20 minutes (1 mark).**

**TEE 2001 Extended Answer:**

Phosphoric acid, $H_3PO_4$, is one of the most widely produced industrial chemicals in the world.

In one method for producing phosphoric acid, $Ca_3(PO_4)_2$ from phosphate rock is heated in an electric furnace with $SiO_2$ and C(graphite). This produces $P_4$ as a hot vapour according to the following equation:

$$2\ Ca_3(PO_4)_2(s) + 6\ SiO_2(s) + 10\ C(s) \rightarrow 6\ CaSiO_3(\ell) + 10\ CO(g) + P_4(g) \qquad \Delta H = -3060\ kJ$$

The reaction mixture can reach a temperature of about 2000 °C. In some modern plants the heat generated is used to power steam turbines.

The $P_4$ is converted into $P_4O_{10}$ by the combustion of phosphorus vapour as it is produced by the furnace ($\Delta H$ = -3053 kJ per mole of $P_4$). This reaction can occur spontaneously at room temperature.

The $H_3PO_4$ is formed by passing the $P_4O_{10}$ through a spray of water in a tower ($\Delta H$ = -377 kJ per mole of $P_4O_{10}$).

Compare and contrast this information about phosphoric acid with the manufacture of sulfuric acid.

**SAMPLE ANSWER:**

**The source of phosphorous for the production of $H_3PO_4$ is phosphate rock ($Ca_3(PO_4)_2$). Sulfur in the contact process can either be sourced directly from solid sulfur, or from sulfur contained inside sulfide ores like pyrite ($FeS_2$). Both of these sources can be oxidised to form sulfur dioxide, as shown in the equations below:**

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

$$4\ FeS_2(s) + 11\ O_2(g) \rightarrow 2\ Fe_2O_3(s) + 8\ SO_2(g)$$

**All of these reactions are exothermic. The reactions for producing $SO_2(g)$ only need oxygen as an additional reactant, which is readily available from the air.**

**In the production of phosphoric acid the next step is to combust the $P_4$ to form $P_4O_{10}$ according to the following equation:**

$$P_4(g) + 5\ O_2(g) \rightarrow P_4O_{10}(g) \qquad \Delta H = -3053\ kJ\ mol^{-1}$$

**A similar step occurs in the sulfuric acid process, where the sulfur dioxide reacts with oxygen to form $SO_3$:**

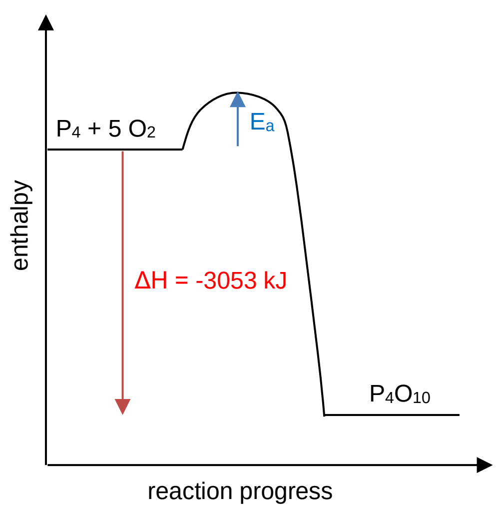
$$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g) \qquad \Delta H = -196\ kJ\ mol^{-1}$$

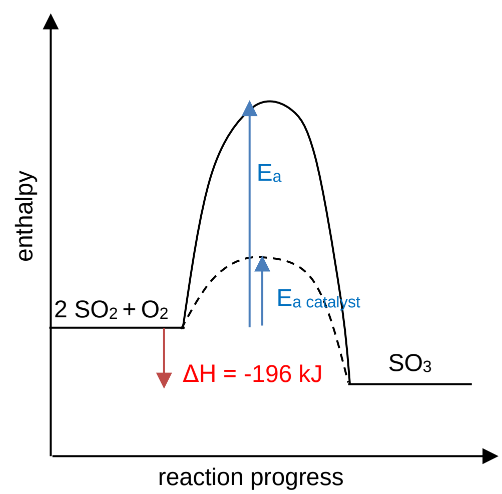
**Both reactions are exothermic, however the reactions involving $P_4$ is much more exothermic than the reaction involving $SO_2$.**

**The conditions for the above listed step differ. The reaction involving $SO_2$ needs moderate temperatures (400-450 °C) whereas the reaction involving $P_4$ can occur spontaneously at room temperature. In both reactions high temperatures would increase reaction rate due to particles having more kinetic energy, thus a greater proportion of particles have energy > $E_a$. The $SO_2$ reaction is reversible, and high temperatures would decrease the yield of $SO_3$ produced because high temperatures would favour the endothermic (reverse) reaction. The moderate temperature used in the production of $SO_3$ is a compromise between acceptable rate and yield. The fact that the reaction involving $P_4$ occurs spontaneously at room temperature might indicate that this reaction has a much lower activation energy than the reaction involving $SO_3$.**

**Another difference between the two reactions is that the oxidation of $SO_2$ involves the use of a catalyst ($V_2O_5$). Catalysts increase reaction rates providing an alternate reaction pathway. They lower activation energy, meaning a greater proportion of particles at a given temperature have $E_k$ > $E_a$. The $P_4$ reaction does not require a catalyst, which could also indicate that it has a lower activation energy than the $SO_2$ reaction.**

**The enthalpy and activation energy differences between the $P_4$ combustion and $SO_2$ combustion reactions can be illustrated through energy profile diagrams:**



**The $P_4O_{10}$ is then reacted directly with water to form phosphoric acid.**

$$P_4O_{10}(g) + 6\ H_2O(\ell) \rightarrow 4\ H_3PO_4(aq) \quad \Delta H = -377\ kJ\ mol^{-1}$$

**The water is sprayed through the gas to maximise its surface area. This reaction can only occur when particles of $P_4O_{10}$ collide with particles of $H_2O$. Increasing the surface area of $H_2O$ increases the area across which collisions can occur, increasing the reaction rate.**

**In comparison, $SO_3$ needs to react in a two-step process to form $H_2SO_4$.**

$$SO_3(g) + H_2SO_4(\ell) \rightarrow H_2S_2O_7(\ell)$$

$$H_2S_2O_7(\ell) + H_2O(\ell) \rightarrow 2\ H_2SO_4(\ell)$$

**The $SO_3$ cannot be dissolved directly in water because the reaction is too volatile and would create an uncontrollable fog of sulfuric acid. This indicates that the reaction between $SO_3$ and $H_2O$ is more exothermic than the reaction between $P_4O_{10}$ and $H_2O$.**

**VCE 2013 Q4:**

The industrial production of hydrogen involves the following two reactions.

Reaction I: $CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$ $\Delta H = +206$ kJ mol$^{-1}$

Reaction II: $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$ $\Delta H = -41$ kJ mol$^{-1}$

(a) Write 'increase', 'decrease' or 'no change' in the table below to identify the expected effect of each change to reaction I and reaction II on the equilibrium yield of hydrogen. (3 marks)

| **Change to reaction I and reaction II** | **Effect of the change on the hydrogen yield in reaction I** | **Effect of the change to the hydrogen yield in reaction II** |
|---|---|---|
| addition of steam at a constant volume and temperature | **increase** | **increase** |
| increase in temperature at a constant volume | **increase** | **decrease** |
| addition of a suitable catalyst at a constant volume and temperature | **no change** | **no change** |

(b) Explain the effect of decreasing the volume, at constant temperature, on the hydrogen equilibrium yield in each reaction. (4 marks)

reaction I: **Equilibrium yield of H2 decreases.**

**Decreasing the volume increases the overall pressure. The system moves to partially compensate and decrease the pressure by favouring the side with fewer moles of gas. i.e. the reverse reaction is favoured.**

reaction II: **Equilibrium yield of $H_2$ does not change.**

**Although the pressure increases, the system is not pushed out of equilibrium because there is the same moles of gas particles on both sides of the equation.**