Ref: Examination Study Questions in TEE Chemistry. Nick Lucarelli. (1998) pp 129-147.

UNIT 7

REACTION RATE AND EQUILIBRIUM

MULTIPLE CHOICE

1. Which of the following statements about reaction rates is correct?

   (i) The orientation of colliding particles can affect their ability to react.
   (ii) The initial speed of a reaction depends upon the enthalpy change, ΔH of the reaction.
   (iii) Reacting particles must collide with sufficient energy if they are to react.

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

2. Powdered aluminium readily burns when sprinkled through a Bunsen flame yet the same amount of aluminium in strip form is difficult to ignite when heated in the same Bunsen flame. Which of the following gives the **best** basis for an explanation of the differing reaction rate?

   (a) collision energy
   (b) concentration of aluminium
   (c) concentration of oxygen
   (d) rate of collision
   (e) orientation

3. Zinc reacts much more rapidly with hot concentrated hydrochloric acid than cold concentrated hydrochloric acid. Which of the following alternatives gives the **best** explanation for the faster reaction rate?

   (a) increased rate of collision
   (b) decreased activation energy
   (c) alternative reaction pathway
   (d) heat of reaction
   (e) increased collision energy

4. Which of the substances listed is used as a **catalyst** in the matching industrial process?

	Industrial process	**Substance**
(a)	Haber process for synthesis of ammonia	MnO_2
(b)	Contact process for synthesis of sulfuric acid	V_2O_5
(c)	Hall-Heroult process for extraction of aluminium	Na_3AlF_6 (cryolite)
(d)	Macarthur-Forrest process for extraction of gold	NaCN
(e)	Blast furnace extraction of iron	$CaCO_3$ (limestone)

5. In the industrial preparation of ammonia by the Haber process the temperature is raised to around 500 °C. Which of the following correctly states the effect of using a high temperature?

   (i) increases the equilibrium yield of NH_3
   (ii) increases the rate of formation of NH_3
   (iii) favours the forward reaction forming NH_3

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

6. The following reaction occurs in a sealed flask.

   $$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l) \qquad \Delta H = -14 \text{ kJ}$$

   Which of the following will increase the **rate** of this reaction?

   (a) cooling the vessel
   (b) increasing the pressure
   (c) decreasing the pressure
   (d) adding some base
   (e) using more finely divided $CaCO_3(s)$

7. Which of the following statements about reaction rates is **false**?

   (a) Increasing the concentration of reactants increases the reaction rate.
   (b) The rate of a reaction involving a solid can be increased by finely dividing the solid.
   (c) Exothermic reactions proceed at a slower rate if the temperature is increased.
   (d) As a reaction proceeds it may slow down due to decreasing concentration of reactants.
   (e) Increasing pressure on a gaseous reaction mixture will increase the reaction rate.

8. Consider this reaction:

   $$2C(s) + O_2(g) \rightarrow 2CO(g) \qquad \Delta H = -222 \text{ kJ}$$

   Which of the following will increase the rate of this reaction?

   (a) lowering the temperature
   (b) lowering the partial pressure of $CO(g)$
   (c) increasing the partial pressure of $O_2(g)$
   (d) lowering the surface area of carbon
   (e) lowering the total pressure

9. The process shown here has reached equilibrium in a closed container at room temperature.

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

Which of the following statements concerning this system at equilibrium is **true**?

(a) Changing the pressure will not affect the equilibrium concentrations.
(b) Adding a catalyst will favour the products and thus lower the pH.
(c) Adding a catalyst will favour the products and thus raise the pH.
(d) Adding $NaOH(s)$ would have no effect on the equilibrium concentrations.
(e) The $Cr_2O_7^{2-}(aq)$ ion is neither being consumed nor produced.

10. This reversible reaction has reached equilibrium:

$$Co^{2+}(aq) + 6NH_3(aq) \rightleftarrows Co(NH_3)_6^{2+}(aq)$$

A small sample of $CoCl_2(s)$ was then added. As it dissolved the concentration of $Co^{2+}(aq)$ initially increased. When equilibrium is again established, how will the concentration of all species compare to their original concentration, ie **prior to addition of** $CoCl_2(s)$?

	$Co^{2+}(aq)$	$NH_3(aq)$	$Co(NH_3)_6^{2+}(aq)$
(a)	same	lower	higher
(b)	higher	lower	higher
(c)	higher	higher	lower
(d)	lower	lower	higher
(e)	same	higher	lower

11. A mixture of nitrogen, hydrogen and ammonia is present in a sealed container at 300 °C and 1020 kPa. The mixture has reached equilibrium according to the following equation:

$$N_2(g) + 3H_2(g) \rightleftarrows 2NH_3(g)$$

The system is then compressed (at constant temperature) to half its original volume. Which of the following correctly describes the changes in **molar amounts** of N_2, H_2 and NH_3 once equilibrium has been re-established?

	N_2	H_2	NH_3
(a)	decreased	decreased	increased
(b)	increased	increased	increased
(c)	increased	increased	decreased
(d)	same	same	decreased
(e)	increased	increased	same

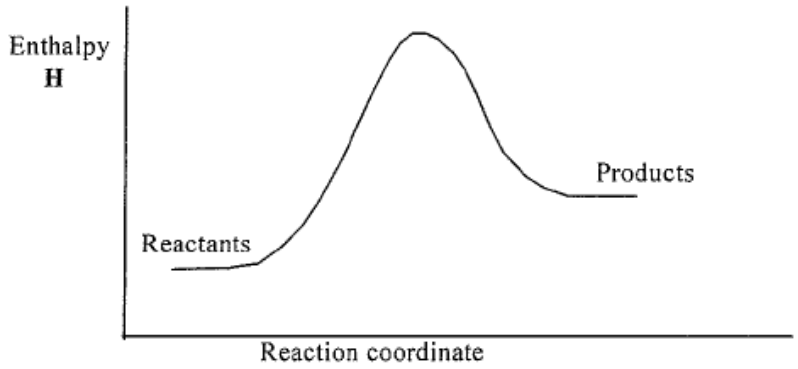
12. To a sealed container is added some $N_2(g)$, $CO_2(g)$ and $H_2O(l)$. After some time the following chemical equilibrium is established:

$$2CO_2(g) + 2H_2O(l) \rightleftarrows 3H^+(aq) + HCO_3^- + CO_3^{2-}(aq)$$

Which of the following changes would **increase** the equilibrium partial pressure of $CO_2(g)$?

(a) removing some of the $N_2(g)$
(b) adding some base
(c) adding a source of $Ca^{2+}(aq)$
(d) increasing the volume of the system
(e) adding some HCl solution

13. A reaction has the following potential energy diagram.



Choose the **false** statement regarding this reaction.

(a) It represents an endothermic reaction.
(b) Enthalpy change for this reaction is positive.
(c) This reaction would cause a temperature rise.
(d) The reverse reaction is exothermic.
(e) Activation energy for the forward reaction is higher than for the reverse reaction.

14. Which of the following is a correct statement about activation energy?

(a) It is equal to the enthalpy change of a reaction.
(b) It is equal to the enthalpy of reactants.
(c) It is equal to the enthalpy of products.
(d) It is equal to the enthalpy of products minus the enthalpy of reactants.
(e) It is equal to the enthalpy of the activated complex (transition state) minus the enthalpy of reactants.

15. A small increase in temperature sometimes produces a large increase in reaction rate. Which of the following statements is the **best** explanation for this?

(a) The reaction must be endothermic.
(b) This increases the rate of collisions between particles.
(c) This causes a lowering of the activation energy.
(d) The transition state becomes less stable.
(e) The colliding particles have higher energy.

16. The reaction shown here proceeds rapidly at room temperature:

$$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$$

Choose the **false** statement regarding this reaction.

(a) The reaction rate increases as the concentration of $H^+(aq)$ decreases.
(b) The reaction rate increases if the zinc is powdered.
(c) The reaction rate may be described in terms of volume of $H_2(g)$ produced per second.
(d) The reaction rate may be described in terms of the mass of $Zn(s)$ consumed per second.
(e) The reaction has a low activation energy.

17. Which of the following statements concerning the activated complex (transition state) of a chemical reaction is **false**?

(a) The activated complex is an unstable substance.
(b) The activated complex always decomposes.
(c) The activated complex may decompose to form products.
(d) The activated complex for the forward and reverse reaction are different.
(e) The activated complex has a higher enthalpy than either reactants or products.

18. One of the reactions involved in the manufacture of sulfuric acid is shown here:

$$2SO_2(g) + O_2(g) \rightleftarrows 2SO_3(g) + \text{Heat}$$

A deliberate effort is made to cool this reaction as it proceeds to equilibrium. Which of the following is the best explanation for doing this?

(i) It increases the equilibrium yield of SO_3 (g).
(ii) It increases the rate of attainment of equilibrium.
(iii) This increases the rate of the forward reaction.

(a) (i) only
(b) (ii) only
(c) (iii) only
(d) (i) and (ii) only
(e) (ii) and (iii)

19. A chemical reaction has reached equilibrium in a closed system. Which of the following observations is **least** likely for this system?

(a) It has a constant colour.
(b) It contains equal amounts of reactants and products.
(c) Reactions in the system will still be occurring.
(d) Its temperature is constant.
(e) It has a constant pressure.

20. Choose the **best** statement about a chemical reaction with an equilibrium constant of 1.

(a) The reaction goes to completion.
(b) Reactants and products have similar concentrations at equilibrium.
(c) At equilibrium there will be mainly reactants.
(d) The reaction is exothermic.
(e) The reaction will quickly reach equilibrium.

21. This reaction has reached equilibrium in a closed system. Which of the following conditions has no effect on the equilibrium system?

$$CaCO_3(s) + 2H^+(aq) \rightleftarrows CO_2(g) + H_2O(l) + Ca^{2+}(aq) + \text{Heat}$$

(a) state of subdivision of the $CaCO_3(s)$
(b) pressure exerted on the system
(c) volume of the system
(d) temperature of the system
(e) concentration of the $H^+(aq)$

22. Which of the following conditions affects the value of the equilibrium constant for a reaction?

(a) temperature
(b) pressure
(c) concentration
(d) catalyst
(e) surface area

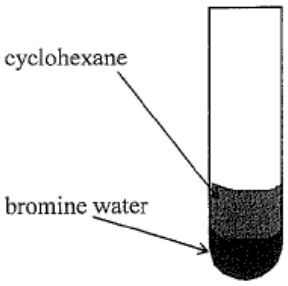
23. Which of the following gives the **best** description of the effect a catalyst has on chemical reactions at equilibrium?

(a) It lowers the value for the equilibrium constant (K).
(b) It favours forward reactions only.
(c) It increases the energy of colliding particles.
(d) It does not affect the rate of reactions in an equilibrium system.
(e) It increases the rate of both forward and reverse reactions.

24. Which of the following set of conditions is closest to that which is used in the manufacture of ammonia by the Haber process?

	Catalyst	Pressure	Temperature
(a)	Yes	3 Atmospheres	50° C
(b)	Yes	300 Atmospheres	500° C
(c)	Yes	3000 Atmospheres	5000° C
(d)	No	3 Atmospheres	50° C
(e)	No	300 Atmospheres	500° C

25. Bromine water and cyclohexene are two immiscible liquids which react to produce 1,2-dibromocyclohexane. The speed of this reaction is increased by shaking the reagents together. Which of the following is the best explanation for the increased rate caused by shaking the reagents?



(a) An improved orientation for colliding molecules.
(b) An increased rate of collision between reacting molecules.
(c) An increase in collision energy.
(d) A catalytic effect due to shaking.
(e) A reduced activation energy for the reaction.

SHORT ANSWER

1. Ethene may be produced from ethane according to the following reversible reaction.

$$C_2H_6(g) \rightleftarrows C_2H_4(g) + H_2(g) \qquad \Delta H = 120\ kJmol^{-1}$$

(a) State three conditions which would increase the rate of the forward reaction.

(i) _____

(ii) _____

(iii) _____

(b) State two conditions which would increase equilibrium yield, ie favour products.

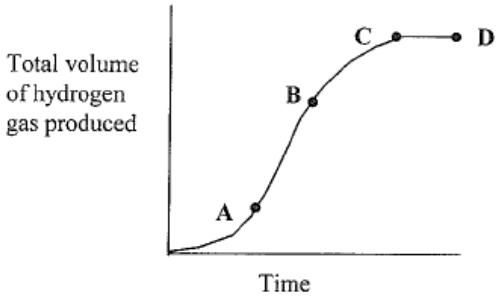
(i) _____

(ii) _____

2. The preparation of hydrogen gas in the laboratory uses the reaction of zinc granules with hydrochloric acid solution.

$$Zn(s) + 2H^+(aq) \rightarrow H_2(g) + Zn^{2+}(aq) \qquad \Delta H = -154 \text{ kJ}$$

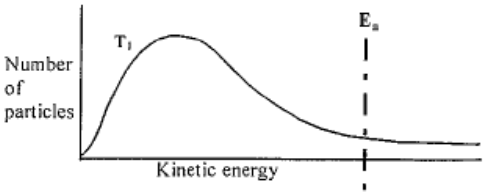
A student investigated the rate of this reaction by preparing a mixture of $HCl(aq)$ and **excess** zinc granules. The total volume of hydrogen produced over time was then measured and recorded. These results were graphed and a sketch of this is shown at right.



Complete the table to describe how the rate of reaction is changing in the various time intervals indicated. Describe them as ***increasing, decreasing, steady or zero***. Also give an explanation for the changing rate in each of the intervals listed.

Interval	**Rate**	**Why the rate of reaction is changing as described**
from the start up to A		
between B and C		
between C and D		

3. The graph at right shows the distribution of kinetic energy for reacting particles at a temperature of T_1. The activation energy for this reaction (Ea) is also indicated.
Sketch onto this graph the distribution of kinetic energy for the same particles at a higher temperature, T_2.
Use your graph to help explain the effect of increasing temperature on reaction rate.



4. One step in the synthesis of nitric acid involves the following reversible reaction:

$$2NO(g) + O_2(g) \rightleftarrows 2NO_2(g) \qquad \Delta H = -114 \text{ kJ}$$

Assume this reaction has reached equilibrium in a closed container at constant temperature and pressure.

(a) What happens (increase, no change, decrease) to the equilibrium yield of $NO_2(g)$ if the following occur?

(i) The volume of the container is increased _____

(ii) More oxygen is added to the container _____

(iii) A suitable catalyst is added _____

(b) What happens initially (increase, no change, decrease) to the rate of the forward reaction if the following changes are made?

(i) The temperature of the container is increased _____

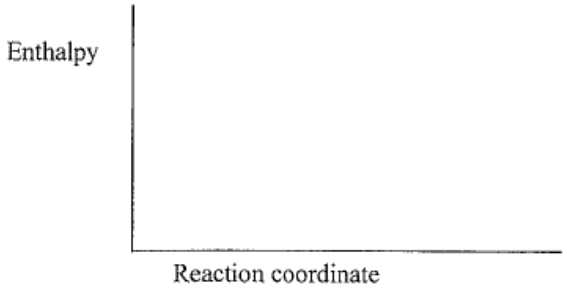
(ii) More $NO_2(g)$ is added into the container _____

(iii) A suitable catalyst is added to the container _____

5. The following equation shows one step in the synthesis of liquid hydrocarbons from methane:

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

Given the reaction is very slow under normal laboratory conditions, sketch an enthalpy change diagram for the reaction. Clearly label ΔH, reactants, products and activation energy.



6. A student prepared a solution of the weak monoprotic acid HF by dissolving 0.100 mole of it in water and making its volume up to 1.00 L in a volumetric flask.

(a) What is the molar concentration of HF*(aq)* molecules if ionisation does not occur?

Answer _____

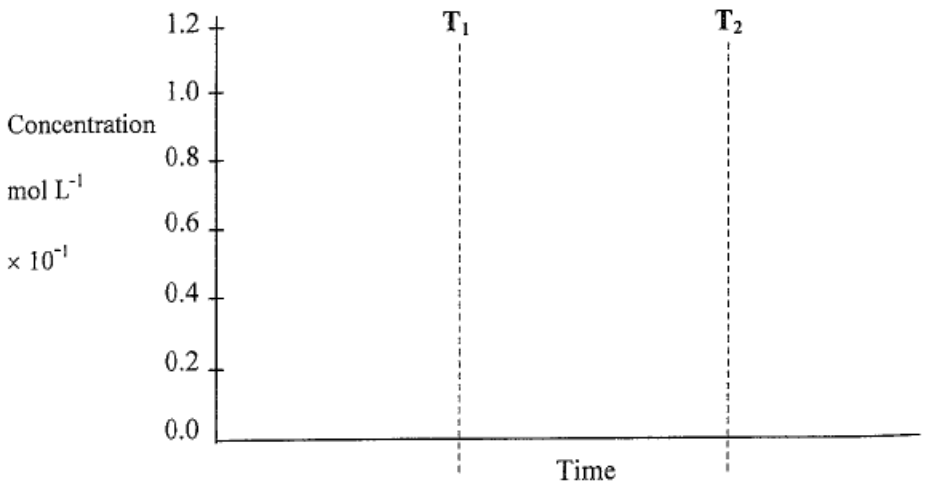
(b) Being a weak acid, **partial** ionisation of HF*(aq)* molecules will occur. Write an equation to represent the formation of $H^+(aq)$ and $F^-(aq)$ by ionisation of HF*(aq)*.

Equation _____

(c) Write an expression for the equilibrium constant for the reaction in (b).

(d) When equilibrium has been achieved the concentration of $H^+(aq)$ and $F^-(aq)$ is 0.025 mol L^{-1} in each case. What will be the actual concentration of HF*(aq)* molecules when equilibrium is achieved? _____

(e) On these axes show the changes in concentration of HF*(aq)*, $H^+(aq)$ and $F^-(aq)$ from when the solution is first made until equilibrium is achieved at time T_1. Ensure your graph is suitably labelled to distinguish HF*(aq)*, $H^+(aq)$ and $F^-(aq)$.

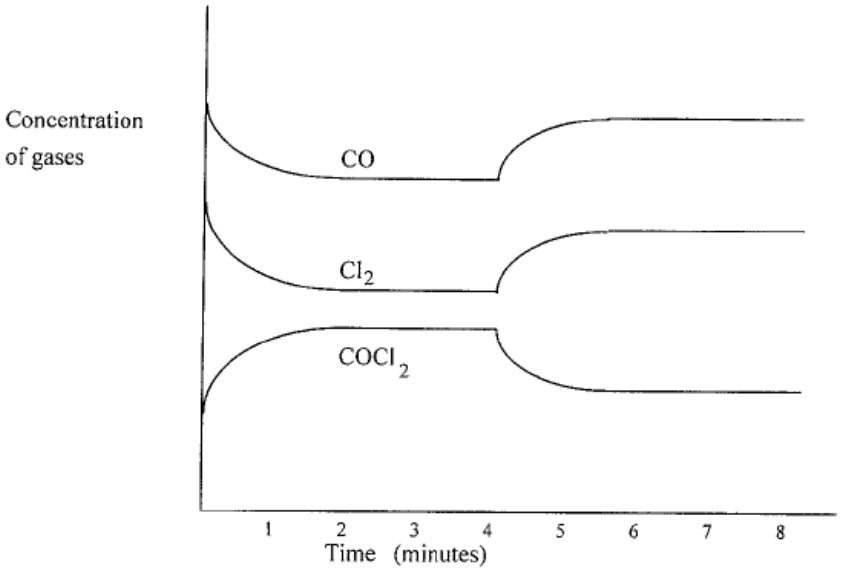


(f) At the time labelled T_2 the student adds some hydrochloric acid to the solution of HF*(aq)* so that the total $H^+(aq)$ concentration rises immediately to 0.05 mol L^{-1}. Continue your graph from T_1 past time T_2 showing how the concentration of the three species HF*(aq)*, $H^+(aq)$ and $F^-(aq)$ changes until a new equilibrium is established. Assume there is negligible volume change on addition of the hydrochloric acid.

7. Phosgene ($COCl_2$) is prepared according to the following reversible reaction:

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

A mixture containing these three gases is introduced into a closed system in the presence of a charcoal catalyst. The following graph shows how the concentration of each of these gases varies with time.



(a) Describe the system three minutes after mixing?

_____

_____

(b) Write an expression for the equilibrium constant of this reaction.

(c) Four minutes after mixing, the temperature of the system is increased to a constant but higher value. From the system's response shown on the graph above, deduce whether the reaction as written is endothermic or exothermic and explain your answer.

Answer _____

Explanation _____

_____

_____

_____

8. Consider this equilibrium process shown below:

$$CO(g) + 2H_2(g) \rightleftarrows CH_3OH(l) + 90\ kJ$$

(a) Write the equilibrium constant expression.

(b) Sketch an enthalpy change diagram for this reaction. Label the axes, reactants, products, activation energy and enthalpy change.

(c) Predict the effect of the following changes on this equilibrium. Write ***favour reactants***, ***favour products*** or ***no change***.

(i) Increasing the pressure on the system _____

(ii) Lowering the temperature _____

(iii) Adding a small amount of $CH_3OH(l)$ (negligible volume change) _____

(iv) Adding a catalyst _____

9. Oxygen can be produced in the laboratory from hydrogen peroxide through the following reaction. Use your knowledge of the collision theory as a basis for explaining the observed effect on reaction rate when the following changes are made to the reaction shown here:

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g) \qquad \Delta H = -948\ kJ$$

(a) When heated with a Bunsen, oxygen is produced much more quickly than when left at room temperature.

_____

_____

_____

_____

(b) Adding a small amount of powdered $MnO_2(s)$ causes the reaction to proceed much more quickly than without it. The $MnO_2(s)$ does not appear to be consumed during the course of the reaction.

(c) Using a more concentrated solution of H_2O_2 causes a more rapid reaction.

10. The economic production of ammonia from nitrogen and hydrogen gas by the Haber process is achieved by applying the principles of equilibrium and reaction rates. Temperature, pressure and catalysts are three conditions which are carefully controlled in this process.

(a) Complete the table below to show your knowledge and understanding of this.

(i) Circle the alternative which comes closest to the conditions of temperature, pressure and catalyst which is used.

(ii) Describe the effect of these conditions on equilibrium ***yield*** of NH_3 and ***rate*** of production of NH_3, ie write ***increase***, ***decrease*** or ***no effect***.

	Conditions used (circle your choice)	**Effect on rate of formation of NH_3**	**Effect on equilibrium yield of NH_3**
temperature degrees Celsius	5 °C 50 °C, 500 °C 5000 °C		
pressure (atmospheres)	3.5 atm 35 atm 350 atm 3500 atm		
catalyst	V_2O_5 Fe/FeO MnO_2 Pt		

(b) Using Le Chatelier's principle explain the effect of pressure on the equilibrium yield of NH_3 in the Haber process.

11. One of the reactions for the manufacture of sulfuric acid by the Contact process is shown here:

$2SO_2(g) + O_2(g) \rightleftarrows 2SO_3(g)$ $\Delta H = -198$ kJ

The following questions are about this reaction.

(a) What is the significance of the two arrows?

_____

_____

_____

(b) The equilibrium constant for this reaction has a very small value at room temperature. What does this indicate about the equilibrium yield of $SO_3(g)$ at room temperature?

_____

_____

(c) In the industrial manufacture of sulfuric acid this reaction is carried out at a temperature of around 450 °C. What is the effect on the equilibrium yield of $SO_3(g)$ and speed of its formation at this higher temperature (compared of normal temperature). Explain your answer in terms of Le Chatelier's principle and the Collision theory.

Effect on equilibrium yield of $SO_3(g)$ _____

Explanation in terms of Le Chatelier's principle _____

_____

_____

_____

_____

_____

_____

Effect on speed of formation of $SO_3(g)$ _____

Explanation in terms of Collision theory _____

_____

_____

_____

_____

EXTENDED ANSWER

When answering the following extended answer questions, use appropriate equations, diagrams and illustrative examples of the chemistry you are describing. Arrange your material as clearly and coherently as possible. Your answer should be presented in about 1½-2 pages.

you will not have to write an essay but you may have to do a one page answer,

1. Write an essay on catalysis, including the following aspects.

   - What are catalysts and how do they effect chemical reactions? Highlight your answer with an example from your laboratory work.
   - Use your knowledge of the collision theory to explain the effect of catalysts.
   - Describe two industrial applications of catalysts.

2. Many chemical reactions involve equilibrium processes. Describe the meaning of "equilibrium" as it applies to chemical reactions. Certain conditions are capable of affecting a system at equilibrium. Note these and describe their effect on chemical equilibrium. Also discuss the application of equilibrium principles in the Haber process (manufacture of ammonia).
   Note: Do not describe the entire Haber process.

ANSWERS - UNIT 7

MULTIPLE CHOICE

1 D	6 E	11 A	16 A	21 A
2 D	7 C	12 E	17 D	22 A
3 E	8 C	13 C	18 A	23 E
4 B	9 A	14 E	19 B	24 B
5 B	10 B	15 E	20 B	25 B

SHORT ANSWER

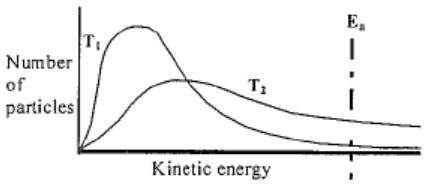
1. (a) (i) - (iii) High temperature, high pressure, catalyst, increase partial pressure of $C_2H_6(g)$.

   (b) (i) - (ii) High temperature, low pressure, reduced partial pressure of $C_2H_4(g)$ or $H_2(g)$.

2.

Start to A	rate increases	This reaction is exothermic thus heat is produced as it progresses. This heat causes an increase in temperature and a corresponding increase in rate as more particles have energy greater than activation energy.
From B to C	rate decreases	The acid (limiting reagent) is almost consumed, thus the concentration of $H^+(aq)$ is low very low. This causes a decrease in the rate of collisions and hence rate decreases.
From C to D	rate is zero	The acid is fully consumed, thus the concentration of $H^+(aq)$ is now zero. Therefore the rate of collisions is zero and hence the rate of reaction is zero.

3.

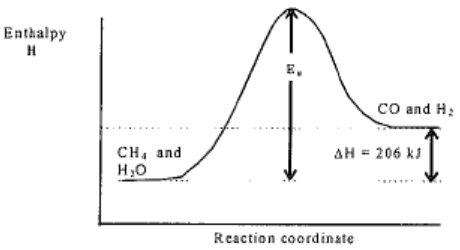


At the higher temperature ($T_2 > T_1$) a greater percentage of particles have a kinetic energy greater than the activation energy (**Ea**). This means a greater percentage of collisions have an energy equal to or greater than the activation energy. Thus a greater percentage of the collisions are successful and hence the reaction rate increases.

4. (a) (i) decrease (ii) increase (iii) no change

   (b) (i) increase (ii) no change (iii) increase

5.



Note: The slow rate of reaction under normal conditions indicates this reaction has a high activation energy.

6. (a) 0.100 mol L^{-1}

   (b) $HF(aq) \rightleftarrows H^+(aq) + F^-(aq)$

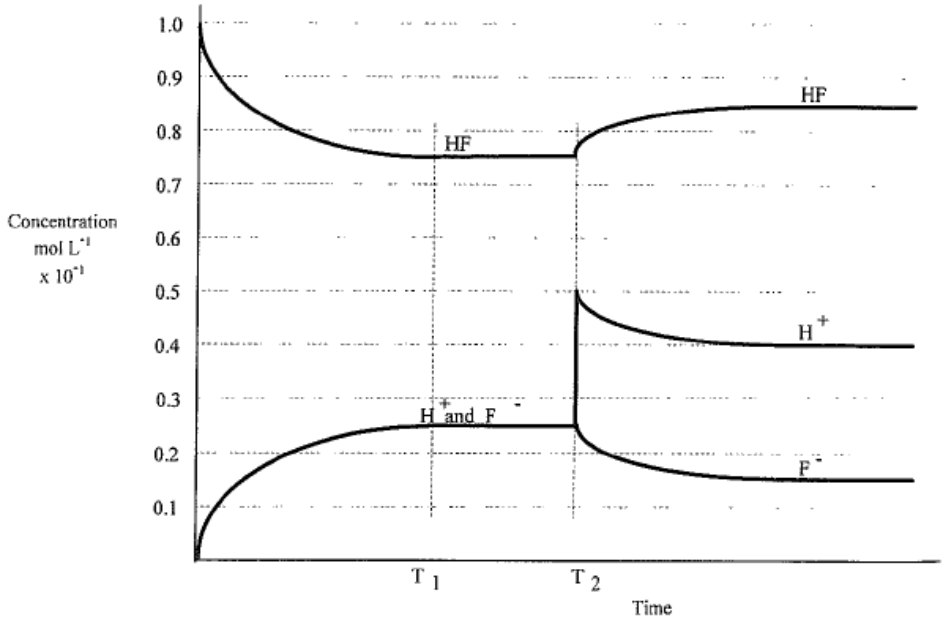
   or

   $HF(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + F^-(aq)$

   (c) $K = \dfrac{[F^-][H^+]}{[HF]}$

   (d) 0.075 mol L^{-1}

(e) and (f) Refer to the graph below.



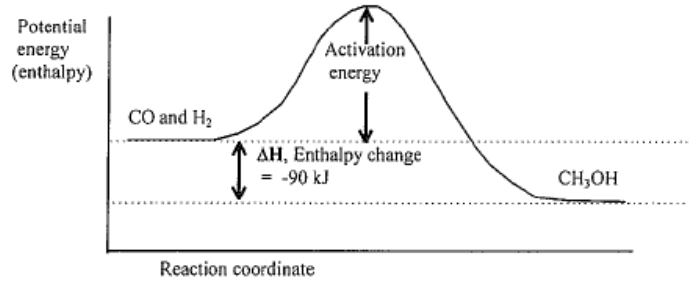
7. (a) The system is at equilibrium. (b) $K = \frac{[COCl_2]}{[CO] \times [Cl_2]}$

(c) Answer: Exothermic

Explanation: The new equilibrium has favoured reactants. Since an increase in temperature favours an endothermic process then the reverse process is endothermic. Hence the forward process is exothermic.

8. (a) $K = \frac{1}{[CO] \times [H_2]^2}$

(b)



(c) (i) favour products (ii) favour products

(iii) no change (iv) no change

9. (a) Using a Bunsen increases the temperature of the reaction mixture. At higher temperatures particles have greater kinetic energy. This means a greater percentage of collisions have an energy equal to or greater than the activation energy. Thus a greater percentage of the collisions are successful and the reaction rate increases producing $O_2(g)$ more quickly.

(b) MnO_2 is a catalyst for the decomposition of H_2O_2. A catalyst provides a reaction pathway with a lower activation energy. As a result a greater percentage of collisions have an energy equal to or greater than the activation energy. Thus more collisions are successful and hence the reaction rate increases.

(c) The higher concentration of H_2O_2 particles causes an increase in the rate of collisions between these particles and hence an increase in the rate of reaction.

10. (a)

Temperature:	500 °C	increases rate	decrease yield
Pressure:	350 atm	increases rate	increases yield
Catalyst:	Fe/FeO	increases rate	no effect on yield

(b) Le Chatelier's principle states, *if a system is at equilibrium and a change in conditions is imposed on the system then the system will re-establish a new equilibrium in such a way as to partially counteract the imposed change*. The formation of ammonia involves a reduction in the total number of moles of gas (4 mol to 2 mol) and therefore a reduction in the total pressure. Thus when the imposed change is an increase in pressure, this system will counteract the change by forming more products, ie increasing the yield of $NH_3(g)$.

11. (a) This shows the formation of $SO_3(g)$ is an equilibrium process. The two arrows represent the forward and reverse reaction of the equilibrium process.

(b) This indicates a low yield of $SO_3(g)$ under normal conditions.

(c) Effect on yield: reduces yield

Explanation: Le Chatelier's principle states, *if a system is at equilibrium and a change in conditions is imposed on the system then the system will re-establish a new equilibrium in such a way as to partially counteract the imposed change*. The imposed change is an increased temperature (450 °C instead of 25 °C) which the system counteracts by favouring the endothermic reaction (this absorbs heat and reduces temperature). For this system the endothermic reaction is the decomposition of SO_3 thus reducing the yield of SO_3.

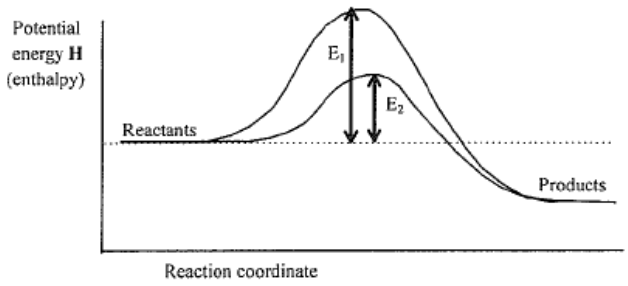
Effect on rate: increases rate

Explanation: At higher temperatures the SO_2 and O_2 particles have greater kinetic energy. This means a greater percentage of collisions between these particles have an energy equal to or greater than the activation energy. Thus a greater percentage of the collisions are successful in producing SO_3 and hence the reaction rate increases. [This effect is even greater for the endothermic reaction and explains the overall decrease in yield of SO_3 described above.]

EXTENDED ANSWER

1. **CATALYSIS**

- Catalysts are substances that increase the rate of a chemical reaction.
- These substances are consumed and produced during the course of a reaction and so there is no net change in the amount of catalyst over time.
- Some examples of catalysts used in the laboratory are:
  - ⇒ $MnO_2(s)$ catalysing the decomposition of H_2O_2
  - ⇒ $H_2SO_4(l)$ acts as a catalyst for esterification reactions.
- For a reaction to occur particles must collide with energy greater than or equal to the activation energy.
- A catalyst enables a reaction to proceed via an alternative pathway of lower activation energy.



E_1 shows the activation energy for the reaction without a catalyst.

E_2 shows the activation energy for the reaction with a catalyst.

- In the presence of a catalyst a greater proportion of collisions have energy greater than or equal to the activation energy. For this reason a greater percentage of the collisions are successful and the reaction rate increases.
- Two industrial applications of catalysts include:

  ⇒ Contact process for the manufacture of H_2SO_4 uses the catalyst V_2O_5 to speed up the complete oxidation of SO_2 gas.

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

  ⇒ Haber process for manufacture of NH_3 uses an Fe/FeO catalyst.

  $$N_2(g) + 3H_2(g) \overset{Fe/FeO}{\rightleftarrows} 2NH_3(g)$$

- In each of the above examples the use of catalysts increases the economic viability of the process by speeding up the relevant reactions.

2. **CHEMICAL EQUILIBRIUM**

Meaning of equilibrium

- Chemical equilibrium involves a balance between two opposing reactions. These reactions are referred to as the forward and reverse reaction.
- At equilibrium the rate of the forward reaction equals the rate of the reverse reaction.
- In a system at equilibrium there are no changes in macroscopic properties such as pressure, temperature, concentration, or colour of the system (ie these properties remain constant).
- Chemical equilibrium can only arise for a chemical reaction occurring in a closed system, that is one where neither matter nor energy can enter or leave the system.

Conditions affecting equilibrium

- The relative concentrations of each species in a system at equilibrium may be altered by changes to:
  - ⇒ the concentration of any one species in the system
  - ⇒ the total pressure on the system
  - ⇒ the temperature of the system.
- **Temperature:** Increasing temperature favours the endothermic process as this consumes heat thus counteracting the increased temperature. Reducing temperature has the opposite effect ie favours the exothermic process.
- Altering temperature has the effect of changing the value of the equilibrium constant.
- **Pressure:** Increasing the total pressure (by reducing volume) favours the side of the reaction with fewest gaseous molecules. By responding this way the system partially counteracts the increased pressure. Decreasing pressure favours the side of the reaction with most gaseous molecules.

- **Concentration:** Increasing the concentration of one species favours the consumption of that species and shifts the equilibrium away from that species. Decreasing the concentration of one species favours the formation of that species and shifts the equilibrium towards that species.
- **Le Chatelier's** principle summarises the effect of these alterations on an equilibrium system, "*if a system is at equilibrium and a change in conditions is imposed on the system then the system will re-establish a new equilibrium in such a way as to partially counteract the imposed change*".

Applying the principles of equilibrium - The Haber process

- Uses high pressure. This favours the formation of fewer gas molecules as this partially counteracts the increased pressure. Thus high pressure favours the formation of $NH_3(g)$ increasing its yield.

$$N_2(g) + 3H_2(g) \rightleftarrows 2NH_3(g)$$
$$\textbf{(4 molecules)} \rightleftarrows \textbf{(2 molecules)}$$

- Uses reduced concentration of NH_3. To achieve this $NH_3(g)$ is frequently removed from the equilibrium mixture. This favours its replacement and forms more $NH_3(g)$ and thus increases the yield of $NH_3(g)$.
- Equilibrium considerations would dictate the use of low temperatures as the forward reaction [formation of $NH_3(g)$] is exothermic. However a moderate temperature is used as this allows a faster reaction rate, resulting in a more economic process despite the reduced yield of $NH_3(g)$.