YEAR 12 CHEMISTRY

Rates & Equilibrium: Answer all questions

Section One: Multiple Choice Questions. (10 marks)

Q1. For the equilibrium

$$2CO_{(g)} + O_{2\,(g)} \rightleftharpoons 2CO_{2\,(g)} \qquad \Delta H = -383 \text{ kJ mol}^{-1}$$

the equilibrium constant would be expected to become larger if

A. the temperature is increased

B. the temperature is decreased

C. the concentration of $CO_{(g)}$ is increased

D. the concentration of $CO_{(g)}$ is decreased

Q2. Iron reacts with steam to set up an equilibrium according to the equation

$$3Fe_{(s)} + 4H_2O_{(g)} \rightleftharpoons Fe_3O_{4\,(s)} + 4H_{2\,(g)}$$

Use of which of the following would **NOT** result in equilibrium being established more quickly?

A. powdered iron instead of lumps of iron

B. a catalyst

C. a larger reaction vessel

D. greater steam pressure

Q3. Ammonia is made industrially by the reaction of nitrogen with hydrogen

$$3H_{2\,(g)} + N_{2\,(g)} \rightleftharpoons 2NH_{3\,(g)}$$

The maximum **YIELD** of ammonia is obtained by ensuring

A. a catalyst is used with as large a surface area as is practicable

B. the equilibrium is established as fast as is practicable

C. the equilibrium lies as far to the right as is practicable

D. the reaction vessel has as big a surface area as is practicable

Q4. A precipitate of sodium chloride could be obtained from a saturated solution of sodium chloride by

A. adding water to it

B. adding concentrated hydrochloric acid to it

C. adding more saturated sodium chloride to it

D. bubbling carbon dioxide solution through it

Q5. In a reversible reaction equilibrium is reached when

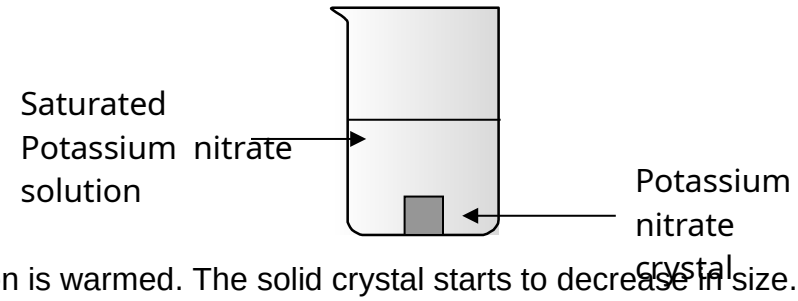
A. reactants stop changing into products

B. the energy of activation of the forward reaction equals that of the reverse

C. the concentration of reactants and products are equal

D. the concentration of reactants and products are constant

Q6. The diagram shows a beaker containing a saturated solution of potassium nitrate and a crystal of solid potassium nitrate.



The solution is warmed. The solid crystal starts to decrease in size.
Which of the following is true about the **concentrations** of solid potassium nitrate and dissolved potassium nitrate.

	Concentration of solid potassium nitrate	Concentration of aqueous potassium nitrate
A	Stays constant	Rises
B	Falls	Rises
C	Rises	Falls
D	Stays constant	Stays constant

Q7. It is important for the pH of blood to be constant within strict limits. One of the equilibria set up in blood to control its pH is

$$CO_{2\,(g)} + H_2O_{\,(l)} \rightleftharpoons HCO_3^-{}_{(aq)} + H^+_{\,(aq)}$$

If the hydrogen ion concentration rises above a certain limit,

A. more carbon dioxide is produced in the blood

B. more hydroxide ions are produced in the blood

C. more carbonate ions are produced in the blood

D. more hydrogen carbonate ions are produced in the blood

Q8. For a reversible reaction **P** $\rightleftharpoons$ **Q**, a rise in temperature will always cause

A. a change in the energy of activation of the forward and reverse reactions

B. an increase in the amount of the products formed

C. an increase in the rates of the forward and reverse reactions

D. a decrease in the rate of either the forward or the reverse reaction

Q9. Each of the reversible reactions below is at equilibrium.

In which one will a change in the volume of the reaction vessel cause **NO** change in the amount of reactants and products, temperature being constant?

A. $H_2O_{(l)} \rightleftharpoons H_2O_{(g)}$

B. $C_{(s)} + O_{2\,(g)} \rightleftharpoons CO_{2\,(g)}$

C. $CO_{2\,(g)} \rightleftharpoons CO_{2\,(aq)}$

D. $3C_2H_{2\,(g)} \rightleftharpoons C_6H_{6\,(g)}$

Q10. The following equilibrium is established in a closed container at constant temperature:

$$2NH_{3\,(g)} \rightleftharpoons 3H_{2\,(g)} + N_{2\,(g)}$$

The addition of the noble gas argon to the equilibrium mixture will cause

A. no change to the amounts of N_2, H_2 and NH_3 gases

B. an increase in the amount of N_2

C. an increase in the amount of NH_3

D. an increase in the amounts of N_2, H_2 and NH_3 gases

End of Section One

Section Two: Short Answer Questions (15 marks)

Q11. This question concerns the reversible reaction

$$3H_{2\,(g)} + N_{2\,(g)} \rightleftharpoons 2NH_{3\,(g)} \; ; \quad K_{eq} = 780 \text{ at } 25\ ^0C$$

This reaction is carried out industrially to manufacture ammonia in the Haber process.

A mixture of nitrogen, hydrogen and ammonia gases in a 10 L vessel at 25 ^{0}C contains 2.0 mol of nitrogen, 2.0 mol of hydrogen and 6.0 mol of ammonia.

A. Has this system reached equilibrium? _____

Give reasons for your choice. _____

_____

_____

_____

_____

B. If the system is not at equilibrium, will the rate of reaction from left to right or the rate of reaction from right to left be the greatest? Explain your answer.

_____

_____

_____

_____

_____

_____

C. Suppose that the mixture has now come to equilibrium at 25 ^{0}C. Analysis shows that the yield of ammonia is smaller than required. Explain why reducing the volume of the vessel at 25 ^{0}C will improve the yield of ammonia at equilibrium.

_____

_____

_____

_____

_____

_____

5 marks

Q12. For the reversible reaction represented by the equation

$$CO_{2\,(g)} + H_{2\,(g)} \rightleftharpoons H_2O_{(g)} + CO_{(g)}$$

A. 2.0 mole of hydrogen gas and 2.0 mole of carbon dioxide gas are mixed in a closed 1.0 litre container and the reaction is allowed to come to equilibrium at 900 K. At equilibrium 0.87 mole of carbon monoxide was present. Calculate the value of the equilibrium constant at 900 K.

_____

_____

_____

_____

_____

_____

B. The following table shows the values of K for the above reaction at various temperatures.

K_{eq}	0.15	0.3	0.95	1.5	2.2	2.9
Temperature K	700	800	1000	1100	1200	1300

Is the forward reaction exothermic or endothermic? Explain

_____

_____

_____

_____

_____

_____

C. If the experiment in A. had been carried out in a 2.0 litre container, what effect would this have on the number of moles of carbon monoxide gas at equilibrium at 900 K? Explain.

_____

_____

_____

_____

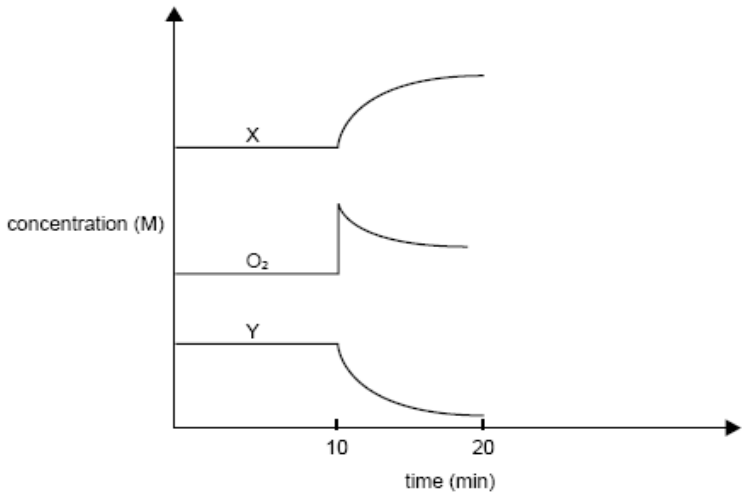
_____5 marks

Q13. 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

$$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)} \qquad \Delta H = -192 \text{ kJ mol}^{-1}$$

As part of a laboratory study of 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 ^{0}C. The container had a fixed volume and was **thermally well insulated**.

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



A. What change occurred at the 10 minute point?

_____

B. Which components of the equilibrium mixture are represented by X and Y?

X: _____ Y: _____

C. Give explanations for the changes in concentration that occur in X, Y and O_2 between 10 and 20 minutes.

_____

_____

_____

D. Would the temperature of the mixture **increase, decrease** or **remain the same** between 10 and 20 minutes? Explain your reasoning.

_____

_____

_____

_____

5 marks

Section Three: Extended Answer Questions (15 marks)

Q14. 200 mL of a solution of HC*l* was partly neutralised by the addition of 1.00 g of pure $CaCO_3$, and 20.0 mL of lime water (containing 1.855 g of $Ca(OH)_2$ per litre) was just sufficient to complete the neutralisation.

What was the concentration (molarity) of the acid solution?

7 marks

15. Outline the collision theory of gaseous reactions. In your discussion, make specific reference to:

A. the significance of the distribution of molecule velocity to reaction rate

B. the effect of temperature on:

- the activation energy

- reaction rate

C. the importance of collision geometry (orientation)

END OF TEST

UNIT3ACHE **SOLUTIONS**

RATES AND EQUILIBRIUM: Answer all questions

Section One: MULTIPLE CHOICE QUESTIONS (10 marks)

1B	2C	3C	4B	5D	6A	7A	8C	9B	10A

Section Two: Short Answer Questions (15 marks)

11.

A. [ ] = molL^{-1}; $[N_2]$ = 0.2 molL^{-1}; $[H_2]$ = 0.2 molL^{-1}; $[NH_3]$ = 0.6 molL^{-1}

$$\text{Reaction Quotient } Q = \frac{[NH_3]^2}{[H_2]^3 \times [N_2]} = \frac{0.6^2}{0.2^3 \times 0.2} = 225$$

This system is not at equilibrium since $Q \neq K$.

B. Since $Q < K$, it follows that the [ ] of the reactants must decrease and the [ ] of the product increase. Therefore the rate of reaction from left to right is greater than the rate of reaction form right to left i.e. equilibrium is being approached from the left.

C. Reducing the volume is equivalent to increasing the pressure. An increase in pressure will favour the production of less moles i.e. the production of ammonia, therefore an improved yield.

12.

A. $[CO] = 0.87$ molL^{-1} $= [H_2O]$

$[CO_2] = [H_2] = 2.0 - 0.87 = 1.13$ molL^{-1}

$$\text{Equilibrium constant } K = \frac{[CO] \times [H_2O]}{[CO_2] \times [H_2]} = \frac{0.87^2}{1.13^2} = 0.593$$

B. Endothermic. As the temperature increases, the value of K increases which indicates increased yield of products. Increasing the temperature shifts the position of equilibrium in a direction so as to partially counteract this change i.e. favours the endothermic process.

C. There would be no change in the number of moles of CO. Since the reaction involves equal numbers of moles of gaseous reactants and equal numbers of moles of gaseous products, decreasing the pressure by increasing the volume would have no effect on the position of equilibrium.

13.

A. $[O_2]$ is increased where [ ] = concentration

B. X: SO_3 Y: SO_2

C. The system is no longer at equilibrium. Since the $[O_2]$ was increased the position of equilibrium will shift in a direction so as to partially counteract this imposed stress, thus producing more SO_3 and reducing the $[O_2]$ and $[SO_2]$ and restoring equilibrium.

D. Increase. The production of SO_3 is exothermic (ΔH = −ve). As more SO_3 is produced the energy produced is released to the surroundings.

Section Three: Extended Answer Questions (15 marks)

14. $n(CaCO_3)$ $= m \div M = 1.00 \div [40.08 + 12.01 + 3(16) = 1.00 \div 100.09 = 0.00999$ mol

$2HCl_{(aq)} + CaCO_{3(s)} = CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

or $2H^+_{(aq)} + CaCO_{3(s)} = Ca^{2+}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

$n(HCl) = 2 \times n(CaCO_3) = 2 \times 0.00999 = 0.01998$

$n(Ca(OH)_2) = m \div M = 1.855 \div [40.08 + 2(2(1.008) + 16)] = 1.855 \div 74.096 = 0.025035$ mol

$n = cV$ therefore $c = n \div V$

$c(Ca(OH)_2) = 0.025035 \div 1$ molL^{-1}

$n(Ca(OH)_2)_{\text{ in 20 mLs}} = 0.025035 \times 0.02 = 0.0005007$ mol

$n(OH^-) = 2 \times n(Ca(OH)_2) = 2 \times 0.0005007 = 0.0010014$ mol

$2HCl_{(aq)} + Ca(OH)_{2(aq)} = CaCl_{2(aq))} + 2H_2O_{(l)}$

or $H^+_{(aq)} + OH^-_{(aq)} = H_2O_{(l)} = 0.0010014$ mol

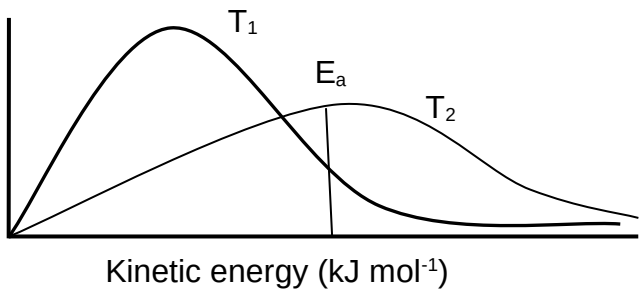
$n(HCl)_{\text{ in total}} = 0.0010014 + 0.01998 = 0.02098$ mol

$n = cV$ therefore $c = n \div V$

$c(HCl) = 0.02098 \div 0.2 = 0.105$ molL^{-1}

15.

A. In any sample of molecules there will be a distribution of velocities, some slow some fast. The majority of molecular velocities will be around the mean (average). Temperature is the mean kinetic energy of the particles ($E_K = \frac{1}{2}mv^2$). Obviously a higher temperature will lead to a higher distribution of faster molecules. A higher distribution of faster molecules will mean a larger number of molecules will have energy ≥ E_a (energy of activation), therefore the reaction rate will be faster at higher temperatures.



B. Temperature has no effect on E_a (Energy of activation) as an increase in temperature only increases the mean E_K (kinetic energy) of the particles. An increase in average E_K energises the particles leading to an increase in the number of particles with energy ≥ E_a (energy of activation). For this reason a rise in temperature increases reaction rate.

C. Many molecules due to their shape or size have active sites in certain areas e.g. the –OH group in alcohols is the active site (region) of the molecule. It follows that collisions in these area (i.e. a fruitful collision) are essential for a reaction to take place