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VCE Chemistry ¾
AOS 2 Revision II [2.10]
Contour Check Solutions

Contour Check

Section A: [2.6] - Rates of Reaction (Checkpoints) (32.5 Marks)

Sub-Section [2.6.1]: Explain How Factors Increase Frequency of Collisions

Question 1 (1.5 marks)

State what happens to the rate of reaction for each of the following scenarios for the following reaction:

$$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(s) + H_2(g)$$

Change	Effect on Rate of Reaction
a. The mass of magnesium metal added is halved.	rate [increases] / [decreases] / [stays same]
b. Magnesium solid is added in clumps.	rate [increases] / [decreases] / [stays same]
c. Concentration of sulphuric acid is decreased to 1.5 M.	rate [increases] / [decreases] / [stays same]

Question 2 (2 marks)

Explain how increasing the volume of the aqueous mixture of the reaction between potassium hydroxide and sulphuric acid will affect the rate of reaction.

By increasing volume, particles drift further apart, resulting in decreased frequency of total collisions (1). This will decrease the frequency of successful collisions with correct orientation, thus decreasing rate of reaction (2).

Space for Personal Notes

Question 3 (2 marks)

Lumps of sodium solid react with hydrochloric acid according to the following equation:

$$2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$$

Explain what would happen to the rate of reaction if finely grounded powder of sodium solid were to be used instead.

If grounded powder of sodium is used instead of lumps, the surface area to volume ratio exposed to HCl would increase (1). This results in an increased frequency of total collisions and hence an increased frequency of successful collisions with correct orientation. This results in an increased rate of reaction. (2)

Question 4 (4 marks)

Eric is preparing two solutions:

Solution 1: Adding 100 ml Silver Nitrate (1.0 M) to 200 ml Sodium Chloride (2.0 M).

Solution 2: Adding 100 ml Silver Nitrate (1.0 M) to 200 ml Sodium Chloride (3.0 M).

Assuming all other conditions are controlled, which solution will have the greater rate of reaction? Explain using Collision Theory.

Solution 2. (1)
Solution 2 contains more concentrated species (1.0 M and 3.0 M), indicating less availability of water molecules in comparison to solution 1 (2). As there is less water volume in solution 2, reactant particles (Ag^+ & Na^+) drift closer together resulting in an increased frequency of total collisions (3). Consequently, the frequency of successful collisions will increase, resulting in an increased rate of reaction (4).
Reactants must be named explicitly for 3rd mark.

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Sub-Section [2.6.2]: Explain How Temperature & Catalyst Affect the Proportion of Successful Collisions

Question 5 (1 mark)

Predict which of the following will have the greatest impact on the rate of reaction when temperature is decreased.

A. The decrease in total number of collisions.

B. The decrease in the mean energy per collision between reactant particles.

C. The gas particles will decrease in size due to a temperature decrease, causing a pressure decrease in the vessel.

D. The reaction vessel shrinking in size due to the exposure to decreased temperatures.

Question 6 (3 marks)

Rehansa sets up a reaction between Sodium iodide (NaI) and hydrochloric acid (HCl) in a reaction vessel. Given that Rehansa decreases the temperature of the reaction vessel, predict the effect on the reaction rate.

By decreasing the temperature of the vessel, the mean kinetic energy of reactant particles will decrease resulting in decreased frequency in total collisions as particles (1). Further, a lower proportion of successful collisions will occur where reactant particles will collide with sufficient force to overcome the activation energy at correct orientation (2). This causes a lower rate of reaction (3).

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Question 7 (1 mark)

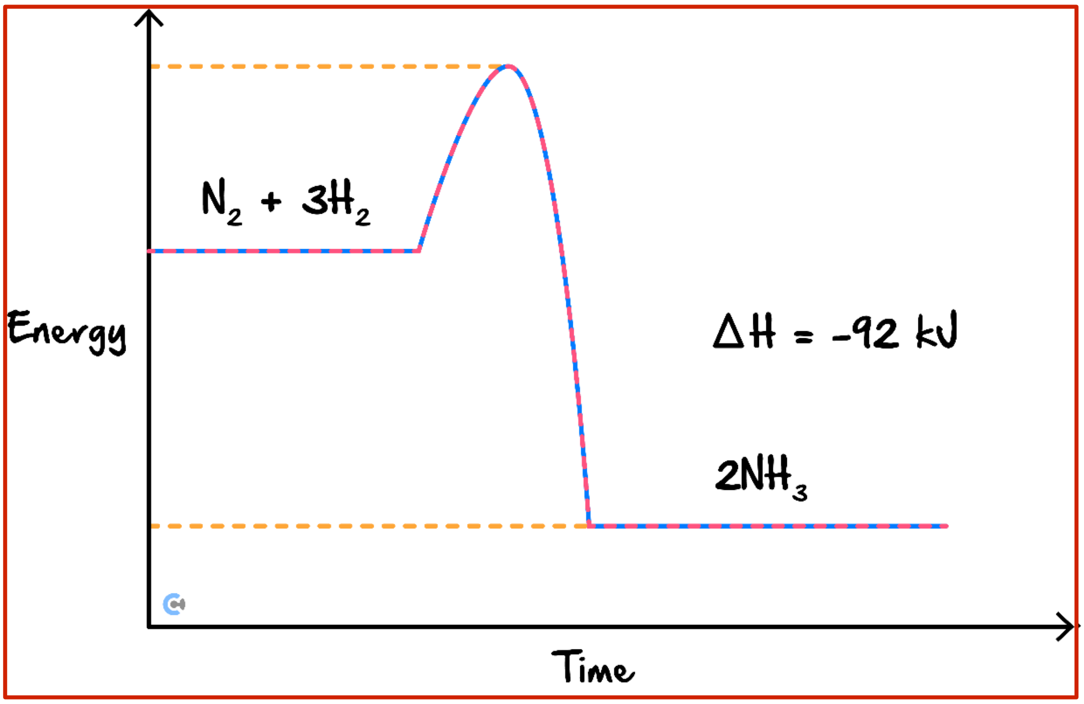
Nitrogen gas is added to a reaction vessel containing hydrogen gas, forming ammonia gas.

Aaliah decreases the temperature in the reaction vessel from 35°C to 10°C.

a. State and justify what happens to the rate of reaction.

By decreasing the temperature of the vessel, the mean kinetic energy of reactant particles will decrease resulting in decreased frequency in total collisions as particles (1). Further, a lower proportion of successful collisions will occur where reactant particles will collide with sufficient force to overcome the activation energy at the correct orientation (2). This causes a lower rate of reaction (3).

b. The energy profile for the original equation is provided before. Draw the new energy profile for the reaction occurring at a lower temperature on the same graph. (1 mark)



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Question 8 (3 marks)

To ensure bread dough rises, it is recommended that the dough is left at room temperature rather than in the fridge. Explain this phenomenon using the collision theory.

At warmer temperatures (such as in the room rather than in the fridge), the mean kinetic energy of particles in the dough (such as yeast and sugars) will be greater, resulting in a greater frequency of collisions (1). Further, a greater proportion of successful collisions will occur where particles collide with sufficient force to overcome the activation energy at correct orientation (2). Hence, there is a greater rate of reaction, meaning that the dough will rise faster (3).

Space for Personal Notes

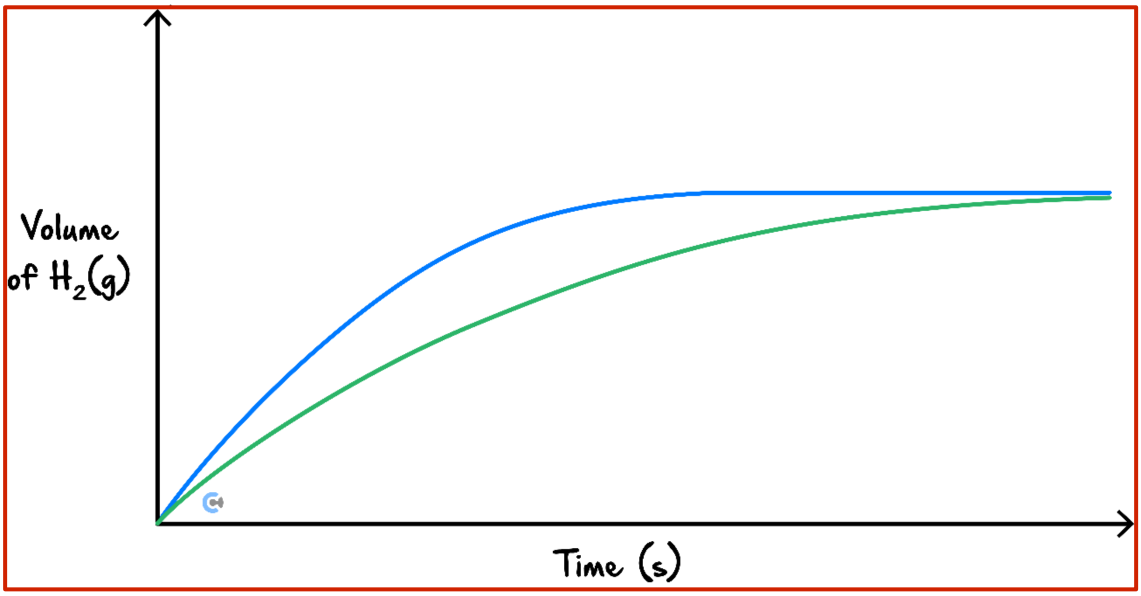
Sub-Section [2.6.3]: Graph Differences in Rate & Yield

Question 9 (3 marks)

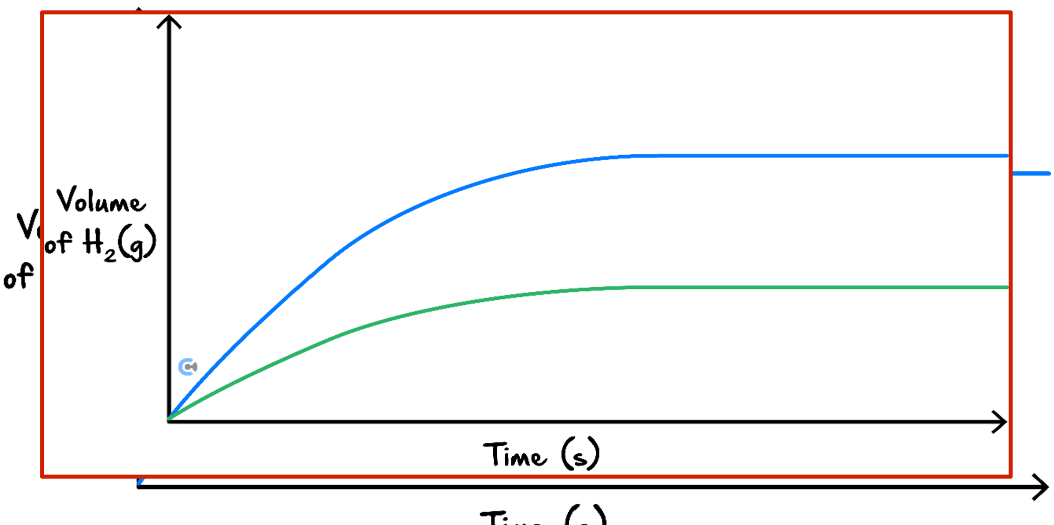
A graph has been provided below, showing how hydrogen gas is formed from a reaction between water and 2.0 g magnesium metal.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

a. Draw with a dotted line, a curve which represents what happens when the temperature is decreased on the graph below. (1 mark)



b. Draw with a solid line, a curve which represents what happens when the mass of magnesium is decreased from 2.0 g to 1.0 g, given that it is the limiting reagent, on the curve below. (2 marks)

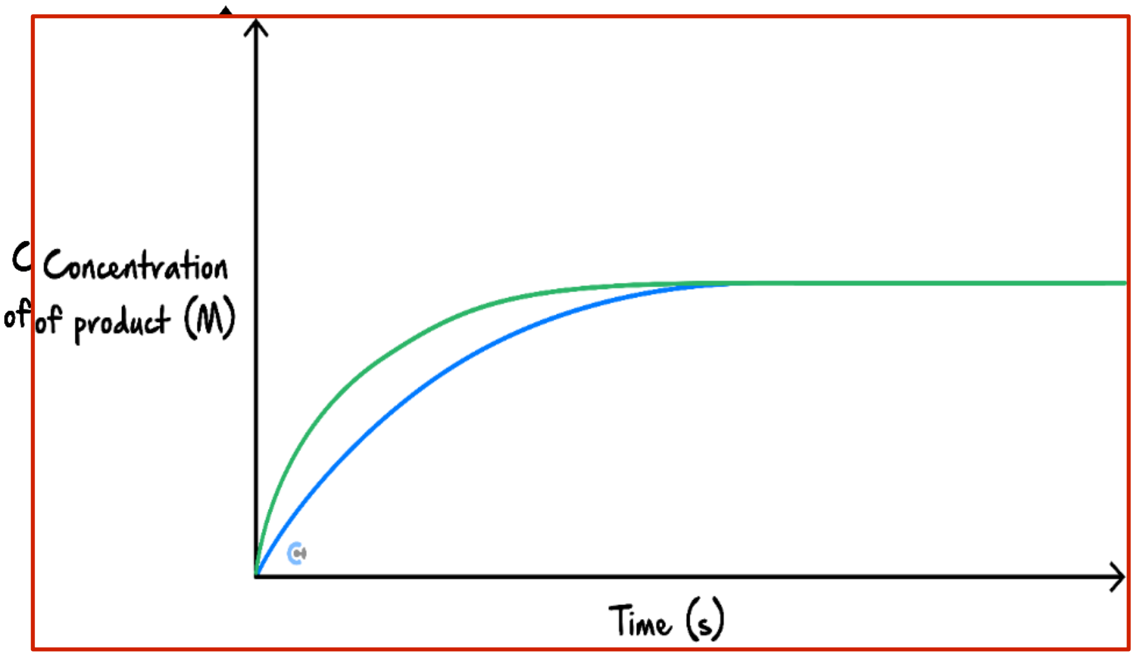


1 mark for decreased yield, 1 mark for the same gradient for the curve.

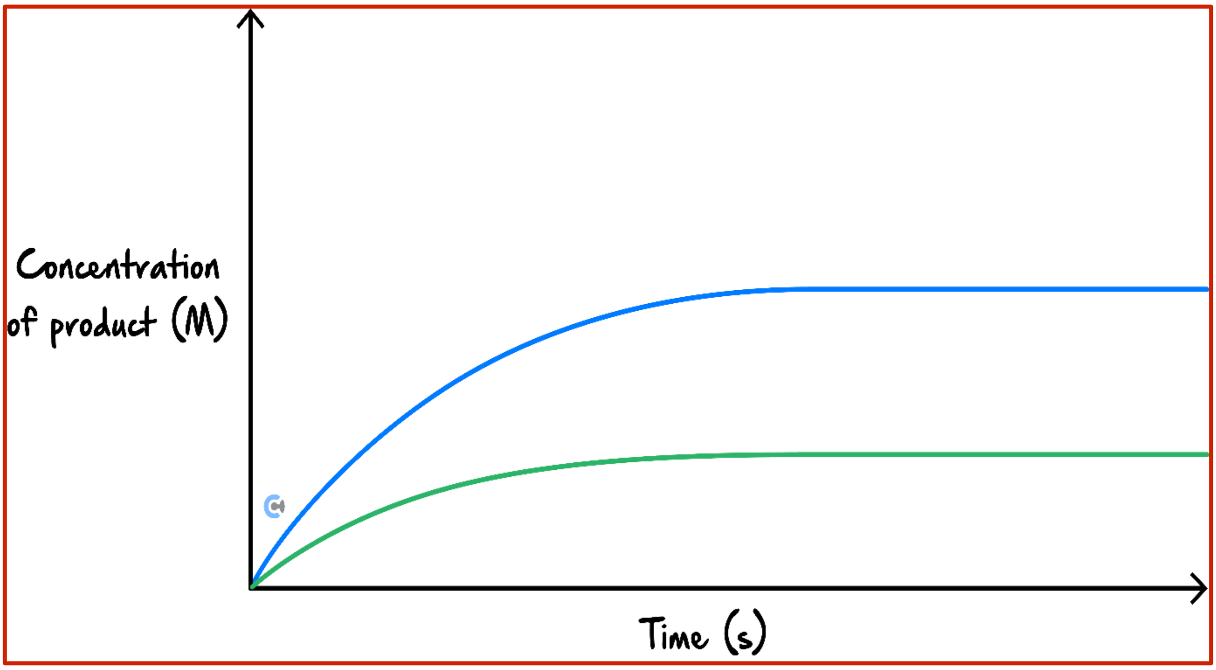
Question 10 (2 marks)

In the graphs below:

a. Draw a curve if the reactant was crushed into a fine powder instead of added as larger pellets. (1 mark)



b. If a reactant was halved in concentration. (1 mark)



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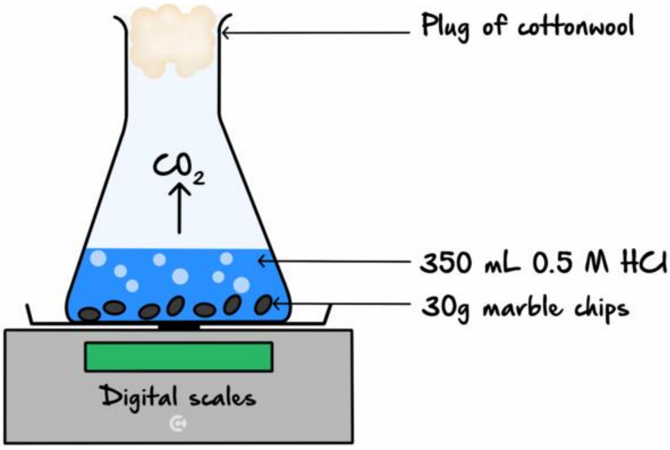
Question 11 (10 marks)

A student set up an experiment to test the effect of different factors on the rate and extent of the reaction between a strong acid and marble chips (Calcium Carbonate, $CaCO_3$). In each trial, the mass of the flask and its contents was measured every 30 seconds, from the instant the reactants were mixed.

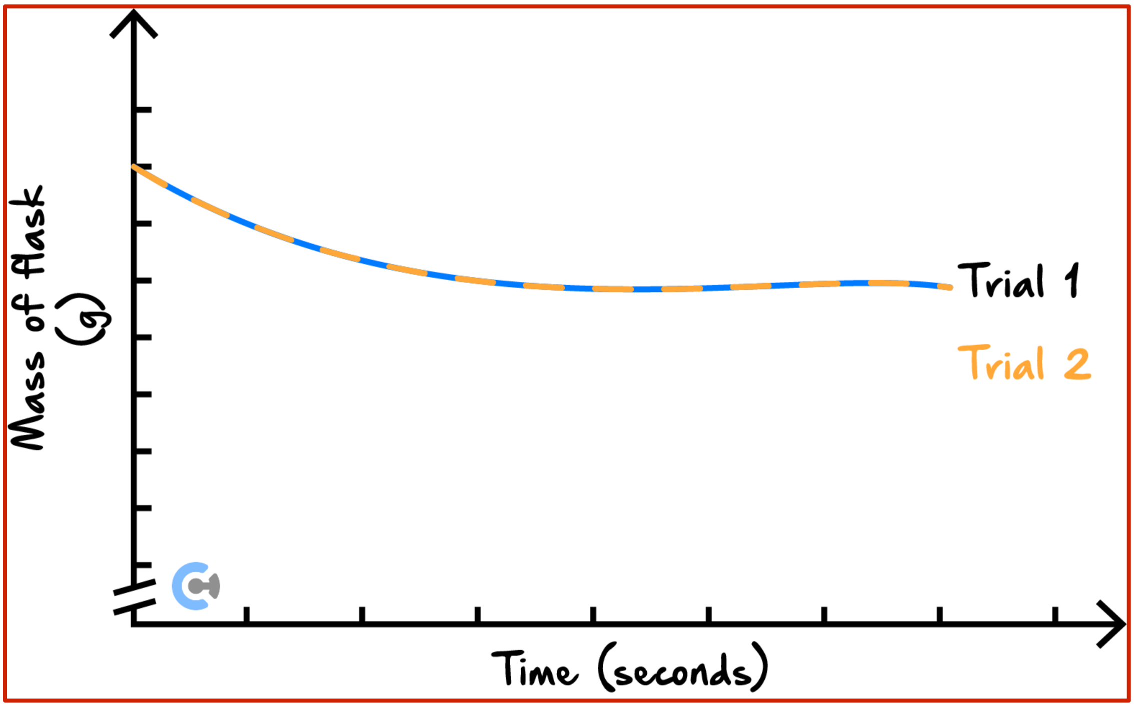
The strong acid used was Hydrochloric acid, HCI. The equation for the reaction is as follows.

$$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

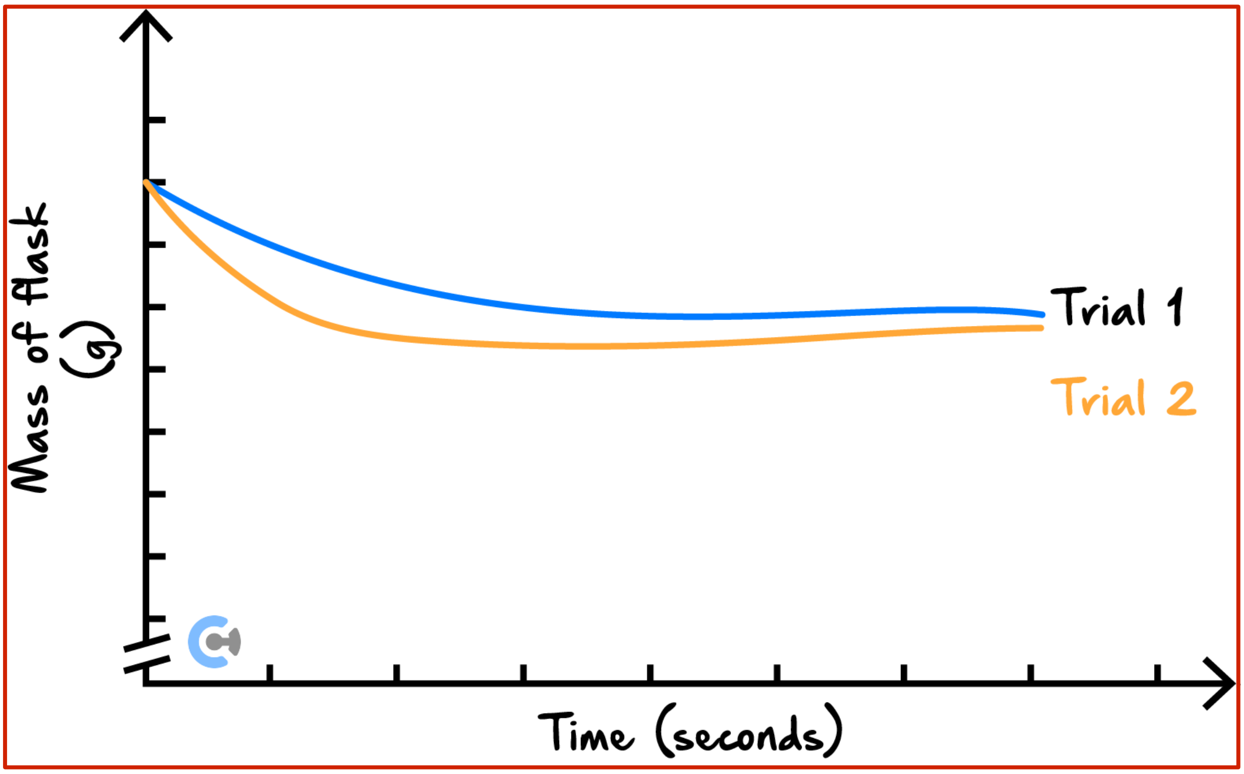


The student then goes about making several changes to the experimental design. Assume trial 1 is the condition of the experiment as shown above.

a. First, the student adds 350 ml of Nitric acid (Trial 2) instead of 350 ml of HCl (Trial 1). On the same set of axes, draw the two trials. (2 marks)

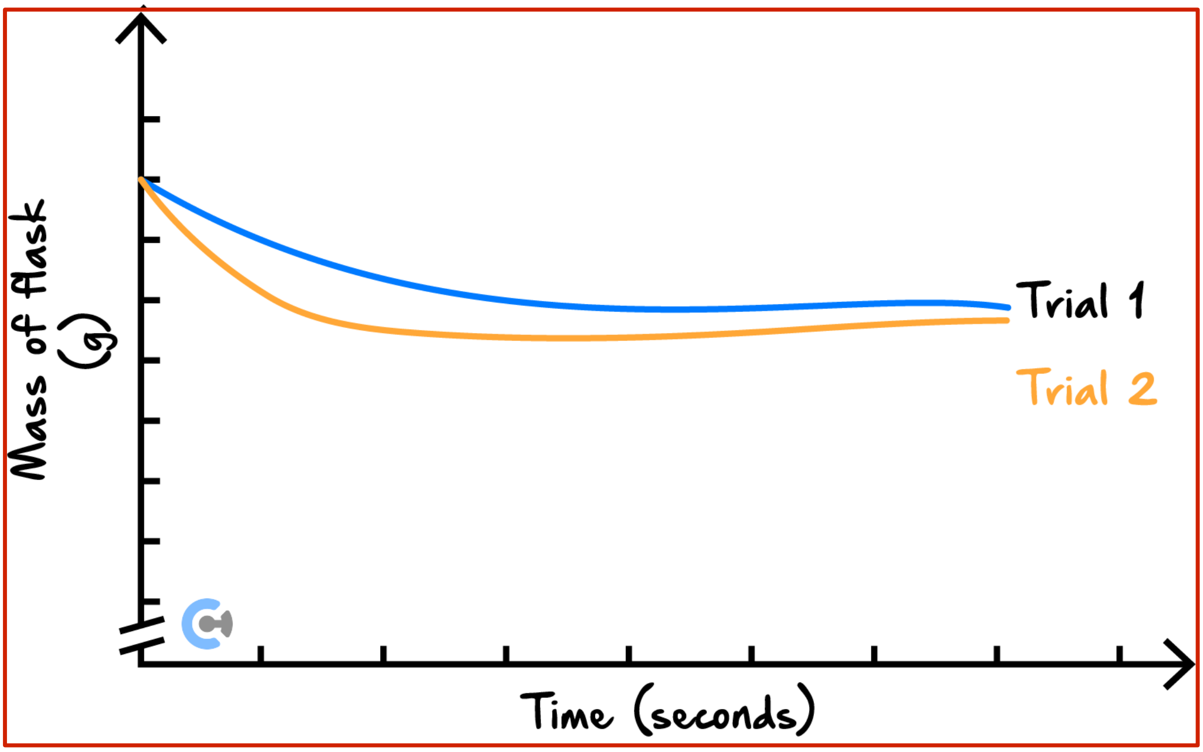


b. The student then crushes the 30 g of marble chips and conducts another trial. Draw what would be expected from the trial on the graph below. (2 marks)



c. The student then decides to add MnO_2 and conducts another trial.

i. Draw what would be expected from the trial below. (2 marks)



ii. Explain why this graph is observed. (4 marks)

MnO_2 provides an alternative chemical reactionary pathway with a decreased activation energy (1). Hence, although the frequency of collisions is the same, a greater proportion of successful collisions occur (2), where reactant Calcium Carbonate and HCl particles will collide with sufficient force to overcome the lower activation energy, thereby increasing the rate of reaction (3). This results in the curve of Trial 2 plateauing quicker than in the absence of a catalyst (4).

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Section B: [2.7] - Equilibrium (Checkpoints) (68 Marks)

Sub-Section [2.7.1]: Write Equilibrium Constant Expression & Find its Value (Including Units)

Question 12 (4 marks)

Medha is interested in the Haber process, for which the chemical equation is shown below:

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

a. In one experiment, at equilibrium, there is 2.15 *mol* of nitrogen gas, 2.5 *mol* of hydrogen gas and 4.55 *mol* of ammonia gas in a 2.0 *L* beaker.

i. Write the K_c expression. (1 mark)

$$K_c = \frac{[NH_3]^2}{[H_2]^3[N_2]}$$

ii. Find the K_c value. (1 mark)

$$K_c = \frac{\left[\frac{4.55}{2}\right]^2}{\left[\frac{2.15}{2}\right]\left[\frac{2.5}{2}\right]^3}$$

$$= 2.5\ M^{-2}$$

b. In another experiment, at equilibrium, Medha finds 3.25 *mol* of nitrogen gas, 1.25 *mol* of hydrogen gas and 3.20 *mol* of ammonia gas in a 4.00 *L* beaker. Find the K_c value.

$$K_c = \frac{[NH_3]^2}{[H_2]^3[N_2]}$$

$$K_c = \frac{\left[\frac{3.2}{4}\right]^2}{\left[\frac{1.25}{4}\right]^3\left[\frac{3.25}{4}\right]}$$

$$= 25.8\ M^{-2}$$

c. Medha's friend, Nawid, is experimenting with the equation shown below:

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

Nawid finds that at equilibrium, 1.05 *mol* of hydrogen gas, 2.10 *mol* of iodine gas and 1.50 *mol* of hydrogen iodide remain in a 3.00 *L* beaker. Find the K_c value. (2 marks)

$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

$$= \frac{\left(\frac{1.5}{3}\right)^2}{\left(\frac{1.05}{3}\right)\left(\frac{2.1}{3}\right)}$$

$= 1.0$ (no units)

Question 13 (4 marks)

Ester hydrolysis is used commercially for methanol production to be used in fuels. One instance of this is the hydrolysis of methyl ethanoate in the presence of water to form ethanoic acid and methanol. The chemical equation is provided below.

$$CH_3COOCH_3(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + CH_3OH(aq)$$

At equilibrium, Hitani finds that 3.5 *mol* of methyl ethanoate, 3.25 *mol* of ethanoic acid and 1.05 *mol* of methanol remains all dissolved in 4.00 *L* of water.

a. Calculate the K_c value. (2 marks)

$$K_c = \frac{[CH_3OH][CH_3COOH]}{[CH_3COOCH_3]}$$

$$= \frac{\left[\frac{3.25}{4}\right]\left[\frac{1.05}{4}\right]}{\left[\frac{3.5}{4}\right]}$$

$= 0.24$ M

b. At a different temperature, Hinati finds that the equilibrium constant is 5.50 M. The 3.125 L vessel contains 1.5 mol of methyl ethanoate, 4.4 mol of water, 2.20 mol of ethanoic acid at equilibrium. Find the concentration of methanol, in M, in the vessel. (2 marks)

$$K_c = \frac{[CH_3OH][CH_3COOH]}{[CH_3COOCH_3]}$$

$$[CH_3OH] = \frac{5.5 \times \left(\frac{1.5}{3.125}\right)}{\left(\frac{2.2}{3.125}\right)}$$

$$= 3.75\,M$$

Question 14 (5 marks)

a. Lachlan is investigating the following reaction:

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

In a 4.5 L vessel at equilibrium, the K_c value is 3.5 M^{-1} and there remains 2.5 mol of $SO_2(g)$ and 1.25 mol of $O_2(g)$ respectively. Find the concentration of sulphur trioxide ($SO_3(g)$). (2 marks)

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

$$[SO_3] = \sqrt{3.5 \times \left[\frac{2.5}{4.5}\right]^2 \times \left[\frac{1.25}{4.5}\right]}$$

$$= 0.55\,M$$

b. Lachlan's friend, Harsh, is interested in the following reaction:

$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$$

In a 10.5 L vessel at equilibrium, the K_c value is 6.3 (no unit) and there remains 1.5 mol of $CO(g)$, 1.25 mol of $H_2O(g)$ and 2.125 mol of $H_2(g)$ respectively. Find the amount, in mol, of carbon dioxide ($CO_2(g)$). (3 marks)

$$K_c = \frac{[H_2][CO_2]}{[CO][H_2O]}$$

$$[CO_2] = \frac{K_c[CO][H_2O]}{[H_2]} = \frac{6.3 \times \left(\frac{1.5}{10.5}\right) \times \left(\frac{1.25}{10.5}\right)}{\left(\frac{2.125}{10.5}\right)} = 0.529$$

$$n(CO_2) = C \times V = 0.529 \times 10.5 = 5.56 \text{ mol}$$

Question 15 (6 marks)

Raph reacts nitrogen gas with hydrogen gas to form ammonia gas.

a. Express the chemical reaction for this scenario. (1 mark)

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

b. Provide the units of K_c for this reaction. (1 mark)

$$K_c = \frac{M^2}{[M^3][M]}$$ $$= M^{-2}$$

c. Write the equilibrium expression for this reaction. (1 mark)

$$K_c = \left(\frac{[NH_{3(g)}]^2}{[H_{2(g)}]^3[N_{2(g)}]}\right)$$

d. In an experiment, 1.20 mol of hydrogen gas, 3.64 mol nitrogen gas and 2.10 mol of ammonia gas was formed at equilibrium in a 2.0 L container at 120°C. Find the equilibrium constant. (3 marks)

$$[NH_3] = \frac{2.1}{2} = 1.05\ M$$

$$[H_2] = \frac{1.2}{2} = 0.6\ M$$

$$[N_2] = \frac{3.64}{2} = 1.82\ M$$

$$\therefore K_c = \left[\frac{[1.05]^2}{[0.6]^3[1.82]}\right]\ M^{-2}$$

$$= 2.804$$

$$\approx 2.8\ M^{-2}$$

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Sub-Section [2.7.2]: Identify the Extent of Reaction

Question 16 (3 marks)

Consider the following chemical reaction:

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

a. If the K_c value is $1.50 \times 10^{6}\ M^{-1}$, state the extent of the reaction. (1 mark)

Large extent of reaction

b. If the K_c value is $9.50 \times 10^{-7}\ M^{-1}$, state the extent of the reaction. (1 mark)

Small extent of reaction

c. If the K_c value is $1.0\ M^{-1}$, state the extent of the reaction. (1 mark)

Medium extent of reaction

Question 17 (3 marks)

Brooke is experimenting with the following reaction.

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^{-}(aq) + H^{+}(aq)$$

At equilibrium in a 10.5 L container, he finds 1.5 mol of CH_3COOH, 1.25 mol of CH_3COO^{-} and 10.5 mol of H^{+} ions.

a. Calculate the K_c value. (2 marks)

$$K_c = \frac{[CH_3COO^{-}][H^{+}]}{[CH_3COOH]}$$

$$= \frac{\left(\frac{1.25}{10.5}\right) \times \left(\frac{10.5}{10.5}\right)}{\left(\frac{1.5}{10.5}\right)}$$

$$= 0.83\ M$$

b. Hence, determine the extent of the reaction. (1 mark)

Medium extent of reaction (1)

Question 18 (1 mark)

When the following reaction reaches equilibrium, it is found that there is a low concentration of nitric dioxide (NO_2) remaining.

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

Hence, predict the magnitude of the equilibrium constant.

The equilibrium constant is likely to be very high ($> 10^4$) due to a high extent of reaction (1).

Question 19 (1 mark)

The value of the equilibrium constant, K_c, for a reaction is 1.0×10^{14}. Which statement about the extent of the reaction is correct?

A. The reaction hardly proceeds.

B. The reaction goes almost to completion.

C. The products have a lower concentration than the reactants.

D. The concentrations of reactants and products are the same.

Space for Personal Notes

Sub-Section [2.7.3]: Find Equilibrium Constant When Equation is Changed

Question 20 (3 marks)

Dai is investigating the following reaction. The equilibrium constant at 25.0°C is 10.6 M^{-1}.

$$NH_3(aq) + H^+(aq) \rightleftharpoons NH_4^+(aq)$$

For the following reactions, state the equilibrium constant at 25.0°C.

Chemical Equation	K_c value at 25°C
a. $2NH_3(aq) + 2H^+(aq) \rightleftharpoons 2NH_4^+(aq)$	$\boxed{112\ M^{-2}}$
b. $\frac{1}{2}NH_3(aq) + \frac{1}{2}H^+(aq) \rightleftharpoons \frac{1}{2}NH_4^+(aq)$	$\boxed{3.26\ M^{-\frac{1}{2}}}$
c. $NH_4^+(aq) \rightleftharpoons NH_3(aq) + H^+(aq)$	$\boxed{0.094\ M}$

Question 21 (3 marks)

Joanne is investigating the following reaction. The equilibrium constant at 35.0°C is 11.6 M^{-1}.

$$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(g)$$

For the following reactions, state the equilibrium constant at 35.0°C.

Chemical Equation	K_c value at 25°C
a. $\frac{1}{2}C_2H_5OH(g) \rightleftharpoons \frac{1}{2}C_2H_4(g) + \frac{1}{2}H_2O(g)$	$\boxed{0.294\ M^{\frac{1}{2}}}$
b. $2C_2H_5OH(g) \rightleftharpoons 2C_2H_4(g) + 2H_2O(g)$	$\boxed{7.4 \times 10^{-3}\ M^2}$
c. $3C_2H_4(g) + 3H_2O(g) \rightleftharpoons 3C_2H_5OH(g)$	$\boxed{1561\ M^{-3}}$

Question 22 (4 marks)

Consider the following equilibrium reaction:

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

a. Write the equilibrium expression for this reaction, including its unit. (1 mark)

$$K_c = \frac{[NO_2]^2}{[N_2O_4]}$$

b. Given that at a certain temperature, the equilibrium constant for this reaction is $K_c = 50\ M$, determine the equilibrium constants for the following modified reactions at the same temperature:

i. $\frac{1}{2}N_2O_4(g) \rightleftharpoons NO_2(g)$. (1 mark)

$$\frac{1}{2}N_2O_4(g) \rightleftharpoons NO_2(g)$$

$$K_c = 50^{\frac{1}{2}} = 7.07\ M^{\frac{1}{2}}$$

ii. $2N_2O_4(g) \rightleftharpoons 4NO_2(g)$. (1 mark)

$$2N_2O_4(g) \rightleftharpoons 4NO_2(g)$$

$$K_c = 50^2 = 2500\ M^2$$

iii. $2NO_2(g) \rightleftharpoons N_2O_4(g)$. (1 mark)

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

$$K_c = \frac{1}{50} = 0.020\ M^{-1}$$

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Sub-Section [2.7.4]: Apply Q_c to Find the Direction of the Equilibrium Shift

Question 23 (3 marks)

Naomi is investigating the following reaction. At 25°C, the system has an equilibrium constant of 35.0 M^2.

$$H_2S(g) \rightleftharpoons 2H^+(aq) + S^{2-}(aq)$$

Determine the direction of equilibrium shift when:

a. $Q_c = 34\ M^2$. (1 mark)

Forward shift

b. $Q_c = 105\ M^2$. (1 mark)

Backward shift

c. $Q_c = 35\ M^2$. (1 mark)

No shift

Space for Personal Notes

Question 24

Claire is interested in the following reaction occurring in a 3.5 L container. She knows that at equilibrium at 26°C, $K_c = 0.1\ M$.

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

She adds 2.0 mol of oxygen gas, 1.5 mol of $SO_2(g)$ and 2.25 mol of $SO_3(g)$. Calculate Q_c and hence predict the direction of the equilibrium shift.

$$Q_c = \frac{[O_2][SO_2]^2}{[SO_3]^2}$$

$$= \frac{\left(\frac{2}{3.5}\right) \times \left(\frac{1.5}{3.5}\right)^2}{\left(\frac{2.25}{3.5}\right)^2}$$

$$= 0.254$$

As $Q_c > K_c$,
system will shift backwards

Question 25 (3 marks)

Jasmine is investigating the following equation occurring in a 10.0 L vessel:

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

Her laboratory technician tells her that at 24°C, $K_c = 30.5\ M$. Jasmine adds 2.0 mol of Fe^{3+}, 10.1 mol of $FeSCN^{2+}$ and an unknown amount of SCN^-. Given that the equilibrium shifts backwards, calculate the SCN^- concentration, in M, which she must have been added.

$$Q_c = \frac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^-]}$$

For equilibrium to shift backward, $Q_c > K_c$

$$30.5 < \frac{\left(\frac{10.1}{10}\right)}{\left(\frac{2}{10}\right) \times [SCN^-]}$$

$$[SCN^-] < \frac{\left(\frac{10.1}{10}\right)}{2/10 \times 30.5}$$

$$[SCN^-] < 0.166\ M$$

Less than 0.166 M of SCN^-

Sub-Section [2.7.5]: Apply RICE Tables to Find K_c

Question 26 (3 marks)

Hayley is running an experiment with the following reaction in a 1.0 L reaction vessel.

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

She adds 1.0 mol of Fe^{3+} and 1.355 mol of SCN^- into an empty reaction vessel. After the system reaches equilibrium, she notes that 0.125 mol of $FeSCN^{2+}$ has been produced. Calculate the K_c value.

	$Fe^{3+}_{(aq)}$	$+ \; SCN^-_{(aq)}$	$\rightleftharpoons FeSCN^{2+}_{(aq)}$
Ratio	1	1	1
Initial	1.0	1.355	0
Change	−0.125	−0.125	+0.125
Equilibrium	0.875	1.23	0.125

$$\therefore \; K_c = \frac{0.125}{1.23 \times 0.875} = 0.116 \; M^{-1}$$

Space for Personal Notes

Question 27 (11 marks)

Clara adds 0.30 mol nitrogen monoxide to 0.40 mol oxygen gas, producing nitrogen dioxide in a 3.0 L vessel in an experiment at school.

a. State the chemical equation for this reaction. (1 mark)

$$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$$

b. Clara conducts the experiment and finds that at equilibrium, 0.165 mol of NO remains.

i. Find the concentration of nitrogen dioxide at equilibrium. (4 marks)

	$2NO_{(g)}$	$O_{2(g)}$	$2NO_{2(g)}$
Ratio	2	1	2
Initial	0.3	0.4	0
Change	0.135	$\frac{0.135}{2} = 0.0675$	0.135
Equilibrium	0.165	0.3325	0.135

$n(NO_2) = 0.135$

$[NO_2] = \frac{0.135}{3} = 0.045$

$\approx 4.5 \times 10^{-2}$ M

ii. Find the equilibrium constant at 15°C. (2 marks)

$[O_2] = \frac{0.3325}{3} = 0.1108$

$[NO] = \frac{0.165}{3} = 0.055$

$K_c = \frac{[0.045]^2}{[0.055]^2[0.1108]}$

$K_c = 6.04$ M^{-1}

c. Another student in Clara's class is conducting the same experiment under the same conditions (15°C, 3.0 L vessel). At one point during the chemical reaction, he notes that there is 0.251 mol of NO, 0.754 mol of O_2 and 1.230 mol of NO_2.

i. Find the reaction quotient. (2 marks)

$$Q_c = \frac{\left[\frac{1.23}{3}\right]^2}{\left[\frac{0.251}{3}\right]^2\left[\frac{0.754}{3}\right]}$$

$$= 95.5\ M^{-1}$$

ii. Comment on the relative rate of production/consumption of reactants and products. Explain how the system will return back to reaching the equilibrium constant. (2 marks)

As $Q_c(95.5\ M^{-1}) > K_c(6.04\ M^{-1})$, at this given point, there is too much products being produced and reactants being consumed (1).
In order to reach the Equilibrium constant, the backwards reaction is favoured in order to produce more reactants and consume more products (2).

Space for Personal Notes

Question 28 (6 marks)

0.500 mol nitrogen gas and 0.400 mol hydrogen gas is added to a 4.0 L vessel, producing ammonia gas.

a. Express the chemical equation for this reaction. (1 mark)

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

b. If the equilibrium constant is known to be 10.65 M^{-2} at 100°C,

i. Express the equilibrium expression. (1 mark)

$$K_c = \frac{[NH_3(g)]^2}{[H_2(g)]^3[N_2(g)]}$$

ii. Find the equilibrium concentration of ammonia gas. (4 marks)

let change of $NH_3(g)$ be represented by x

	$N_2(g)$ +	$3H_2(g)$ $\rightleftharpoons$	$2NH_3(g)$
Ratio	1	3	2
Initial	0.5 mols	0.4 mols	0 mols
Change	$-x$ mols	$-3x$ mols	$+2x$ mols
Equilibrium	$0.5-x$ mols	$0.4-3x$ mols	$2x$ mols

$$k_c = \frac{\left(\frac{2x}{4}\right)^2}{\left(\frac{0.4-3x}{4}\right)^3\left(\frac{0.5-x}{4}\right)} = 10.65$$

$x = 0.0406...$ or $x = 0.900...$

reject $x = 0.900...$ because $0.5 - x > 0$ and $0.4 - 3x > 0$

$\therefore x = 0.0406$

$$[NH_3] = \frac{n}{V} = \frac{2 \times 0.0406}{4.0} = 0.020 \text{ M}$$

Space for Personal Notes

Question 29 (5 marks)

Hydrogen peroxide can be produced from the reversible reaction between hydrogen and oxygen:

a. Write the equilibrium chemical equation. (1 mark)

$$\boxed{H_{2(g)} + O_{2(g)} \rightleftharpoons H_2O_{2(l)}}$$

b. 0.85 *mol* of hydrogen and 0.4 *mol* of oxygen are added to an empty 1.00 *L* reactor. When equilibrium is reached, the amount of hydrogen peroxide present is 0.2 *mol*. Determine the value of K_c. (4 marks)

	$H_{2(g)}$ +	$O_{2(g)}$ ⇌	$H_2O_{2(l)}$
Initial	0.85 mol	0.4 mol	0 mol
Change	−0.2 mol	− 0.2 mol	+0.2 mol
Equilibrium	0.65 mol	0.2 mol	0.2 mol

$$K_c = \frac{[H_2O_2]}{[H_2][O_2]} = \frac{\left(\frac{0.2}{1.00}\right)}{\left(\frac{0.65}{1.00}\right) \times \left(\frac{0.2}{1.00}\right)} = 1.5\ M^{-1}$$

Space for Personal Notes

Section C: [2.8] - Le Chatelier's Principle (Checkpoints) (33.5 Marks)

Sub-Section [2.8.1]: Explain the Effects of Addition/Removal of Substances or Pressure/Volume Changes on Equilibrium System

Question 30 (2.5 marks)

Precious is experimenting with changes made to a sealed reaction vessel containing the following reaction:

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

In the table below, predict the direction in which the equilibrium system will shift (using a left or right arrow, or a dash for no change) when the following changes are made:

S. no	Change	Shift
a.	$N_2(g)$ is removed.	$\boxed{\leftarrow}$
b.	$H_2(g)$ is added.	$\boxed{\rightarrow}$
c.	$NH_3(g)$ is removed.	$\boxed{\rightarrow}$
d.	Volume of the vessel is doubled.	$\boxed{\leftarrow}$
e.	100 ml of water is removed.	$\boxed{-}$

Space for Personal Notes

Question 31 (2 marks)

Melody is interested in the following reaction:

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

Explain the direction in which the equilibrium system will shift if she removes a small amount of ethanoic acid (CH_3COOH).

If ethanoic acid is removed, ethanoic acid (Reactant) concentration will decrease. According to Le Châtelier's Principle, the system will partially oppose the change by increasing the concentration of ethanoic acid (1). Hence, the backward reaction is favoured (2).

Question 32 (3 marks)

Shiven sets up the following reaction in a 2.0 L vessel, which is half full, and allows it to reach equilibrium.

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

Given that Shiven adds 1000 mL of water to the reaction, explain whether the rate of the forward reaction will be greater, equal to, or less than the rate of the reverse reaction.

If 1000 ml of water is added, the volume of the system decreases, increasing the overall concentration of all species will decrease (1). According to Le Châtelier's Principle, the system will partially oppose by increasing the overall concentration back up by favouring the equation with greater particles (2). Hence, the backward reaction is favoured, meaning that the rate of the forward reaction will be LESS than the rate of the backward reaction (3)

Space for Personal Notes

Question 33 (4 marks)

Olivia observes the reaction shown below:

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

She is informed that the $FeSCN^{2+}$ ions produced a distinct red colour whilst all other ions in the system are colourless. At equilibrium, she finds there is 1.5 mol of Fe^{3+}, 1.0 mol of SCN^{-} and 0.75 mol of $FeSCN^{2+}$ present within the 1.0 L vessel.

Olivia adds 0.5 mol of $Fe^{3+}(aq)$ to the system.

When adding iron (III) ions to the solution, she notices that the solution within the vessel increases in intensity to a darker red. Explain this observation.

By adding Fe^{3+} to the system, the concentration of the reactants increases (1). According to Le Chatelier's principle, the system will partially oppose the change by decreasing the concentration of Iron ions (2). This results in the forward reaction being favoured (3), whereby the concentration of $FeSCN^{2+}$ ions increase causing an increase in intensity of the red colour of the solution (4).

Space for Personal Notes

Sub-Section [2.8.2]: Graph Effects of Addition/Removal of Substances or Pressure/Volume Changes on Equilibrium System

Question 34 (3 marks)

Merna allows the following reaction to reach equilibrium.

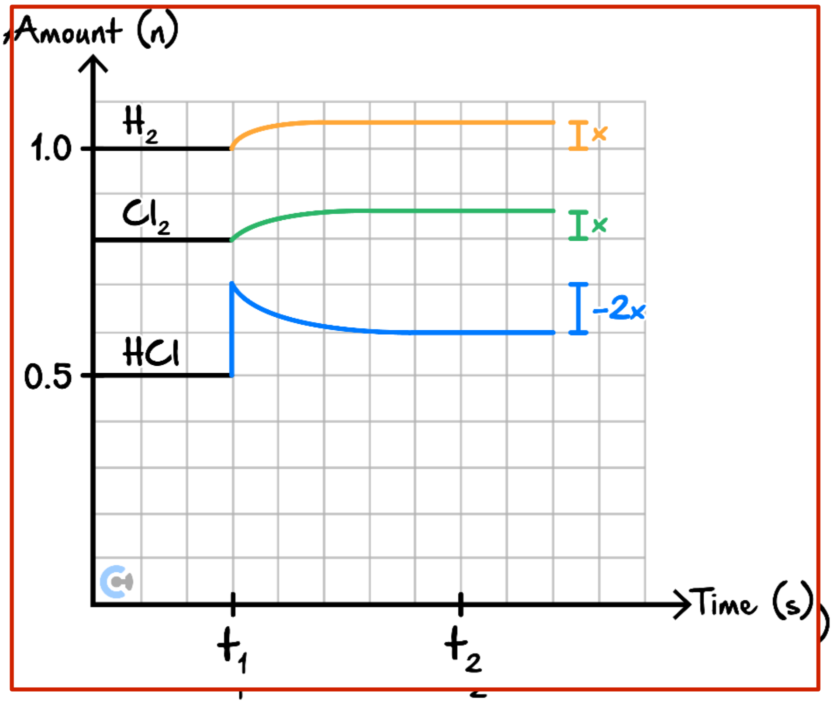
$$H_2(g) + Cl_2(g) \leftrightharpoons 2HCl(g)$$

She adds 0.20 mol of HCl from the vessel.

a. Determine the direction in which the system shifts. (1 mark)

$\boxed{\leftarrow}$

b. Graph how the amounts of the species change with time. Assume HCl is removed at t_1, and equilibrium is re-established at t_2. (2 marks)



Space for Personal Notes

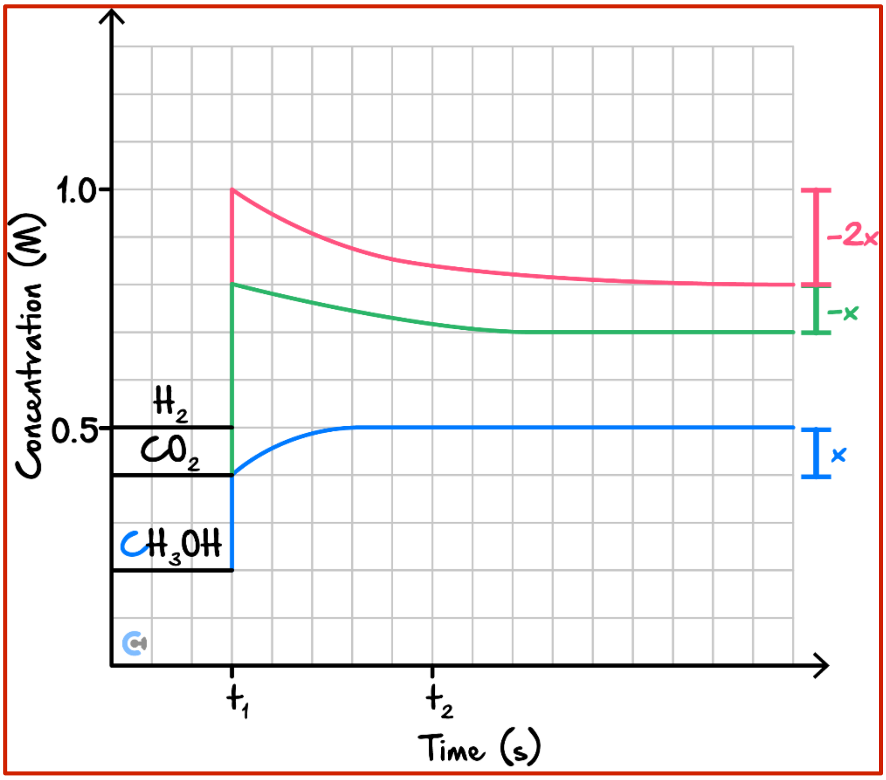
Question 35 (4 marks)

The following reaction is allowed to reach equilibrium in a 2.0 L vessel.

$$CO_2(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$$

Katie halves the volume of the container.

a. Graph how the concentrations of the species change with time. Assume the volume is halved at t_1, and equilibrium is re-established at t_2. (2 marks)



b. When the volume is halved:

i. State the change in the amount of H_2 at t_1. (1 mark)

Stays the same.

ii. State the change in the amount of CO_2 at t_2. (1 mark)

Decreases

Space for Personal Notes

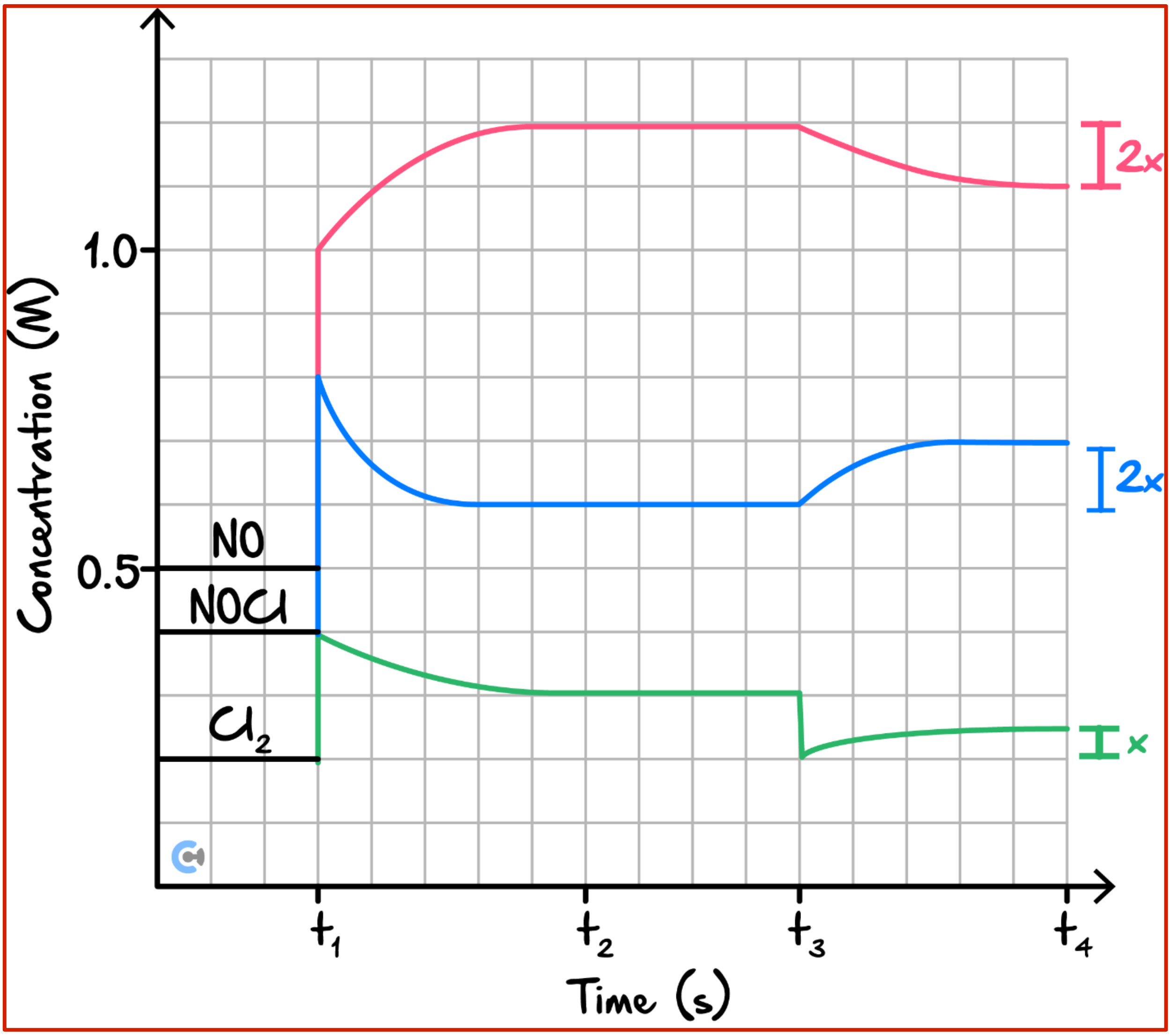
Question 36 (6 marks)

The following reaction is allowed to reach equilibrium in a 3.0 L container.

$$2\text{NO}\,(g) + \text{Cl}_2(g) \rightleftharpoons 2\text{NOCl}(g)$$

Nour halves the volume of the container.

a. Graph how the concentrations of the species change with time. Assume the volume is halved at t_1, and equilibrium is re-established at t_2. (2 marks)



Nour then removes 0.3 mol of chlorine gas (Cl_2) from the system.

b. Graph how the concentration of the species changes with time. Assume the chlorine gas is removed at t_3, and equilibrium is re-established at t_4. (2 marks)

$$\boxed{\left(C = \frac{0.3}{1.5} = 0.2\ M\right)}$$

Sub-Section [2.8.3]: Apply Partial Opposition during Equilibrium to the Effects on Amount, Concentration & Colour of Substance

Question 37 (3 marks)

Claire sets up the following reaction, allowing it to reach equilibrium.

$$SO_2(g) + Br_2(g) \rightleftharpoons SO_2Br_2(g)$$

Claire removes 0.2 mol of Br_2 from the system. Given that the colour of Br_2 is brown, explain what happens to the intensity of the solution when equilibrium is re-established.

When Claire removes Br_2, its concentration decreases, meaning the brown intensity decreases (1). According to Le Chatelier's principle, the system will attempt to partially oppose the change by favouring the backwards reaction (2). This will cause the intensity to increase, but still remain lighter than prior to the removal of the substance as there still is a net decrease in Br_2 concentration (3).

Question 38 (6 marks)

Consider the following reaction where substance B is dark blue.

$$2A(aq) + B(aq) \leftrightarrow C(aq) + D(aq)$$

a. If A is removed, state the change: (2 marks)

i. In the intensity of blue:

Increased Stay the same Decreased

ii. In the amount of A:

Increased Stay the same Decreased

iii. In the amount of C:

Increased Stay the same Decreased

b. If B is added, state the change: (2 marks)

i. In the intensity of blue:

Increased Stay the same Decreased

ii. In the amount of A:

Increased Stay the same Decreased

iii. In the amount of C:

Increased Stay the same Decreased

c. If water is added, state the change: (2 marks)

i. In the intensity of blue:

Increased Stay the same Decreased

ii. In the amount of A:

Increased Stay the same Decreased

iii. In the amount of C:

Increased Stay the same Decreased

Space for Personal Notes

Question 39

The following reaction is allowed to reach equilibrium. $Co^{2+}(aq)$ is pink in colour.

$$Co^{2+}(aq) + 4Cl^{-}(aq) \rightleftharpoons CoCl_4{}^{2-}(aq)$$

Place ticks in appropriate boxes to indicate the effect of each change once a new equilibrium has been established.

Change at Equilibrium	**Colour at New Equilibrium Compared with Initial Equilibrium**		**$[CoCl_4]^{2-}$ at New Equilibrium Compared with Initial Equilibrium**	
	Less Pink	More Pink	Decreased	Increased
Sample 1: 1 drop of $AgNO_3$ solution is added, forming an AgCl precipitate.		✓	✓	
Sample 2: 1 drop of NaCl solution is added.	✓			✓
Sample 3: A large volume of water is removed.		✓		✓
Sample 4: A drop of K_2Co is added.		✓		✓

Space for Personal Notes

Section D: [2.9] - Rate Yield Conflict (Checkpoints) (28 Marks)

Sub-Section [2.9.1]: Explain the Effects of Temperature, Inert Gas, or Catalyst on an Equilibrium System

Question 40 (1 mark)

An inert gas and a catalyst are added to an equilibrium system simultaneously. What is the effect on the equilibrium system?

A. The rates of the forward and reverse reaction remain the same and the concentrations of all reagents increase.

B. The rates of the forward and reverse reactions increase and the concentration of the reagents remains the same.

C. The rate of the forward reaction increases and the concentration of the products increases.

D. Only the rate of the reverse reaction increases and the concentration of the reagents remains the same.

Question 41 (4 marks)

Consider the following equilibrium system.

$$2A(g) + B(g) \rightleftharpoons C(g) \quad \Delta H > 0$$

a. Explain the effect of increasing the temperature of the system. (2 marks)

An increase in temperature will be partially opposed by the endothermic forward reaction (1). This will increase the concentration of the product C. (2)

b. Explain the effect of adding a catalytic powder to the system while increasing the temperature. (2 marks)

Adding a catalyst increases the rate of both the forward and reverse reactions (1) by providing an alternate reaction pathway with a lower activation energy. As a result, the system will establish equilibrium faster after the increase in temperature (2).

Question 42 (4 marks)

The following equilibrium system is used in ColdGlow glow sticks.

$$2X(g) \rightleftharpoons Z(g) + Y(g)$$

ColdGlow glow sticks get brighter based on how cold their surroundings are. Both Z and Y are colourless gases, whereas X has a distinct fluorescent red colour.

a. State whether the above reaction is endothermic or exothermic. Justify your answer. (2 marks)

As the system gets colder (as temperature decreases), the concentration of X increases; the reverse reaction will partially oppose a decrease in temperature and is exothermic (1). The forward reaction is, therefore, endothermic (2).

b. The volume of the glow sticks is decreased and the temperature can be assumed to be constant. Describe and explain any observable effects on the system. (2 marks)

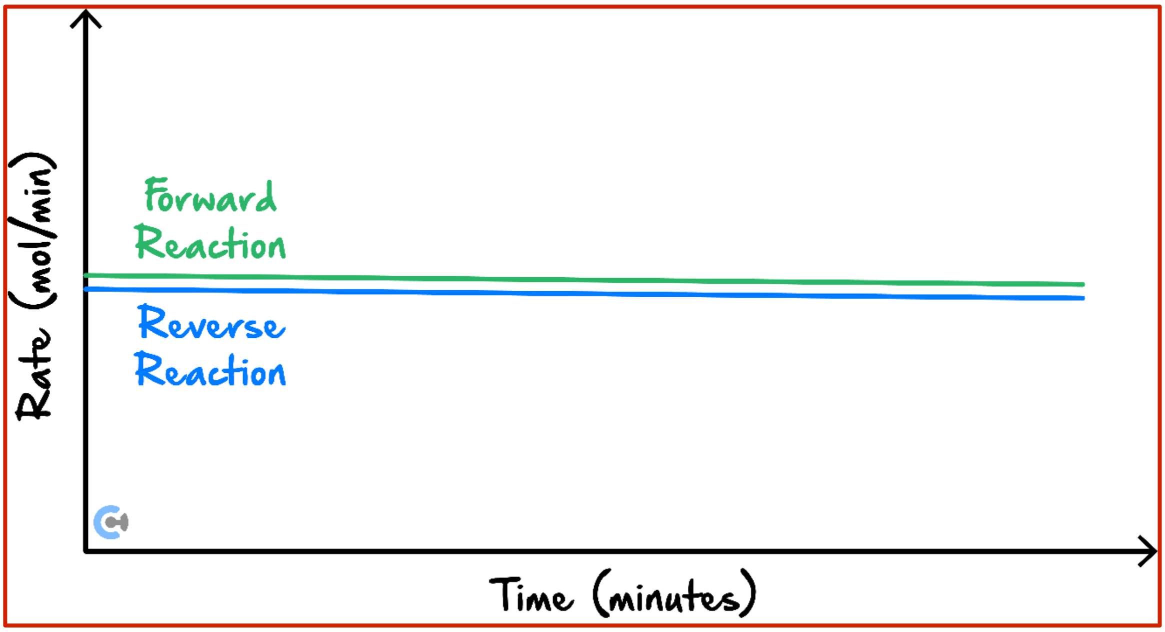
The concentration of all reagents will increase (1) and, as the number of particles on each side of the equation is the same, the concentrations will not decrease. Since the brightness of the ColdGlow glow sticks are directly proportional to the concentration of X, the glow sticks will glow brighter (2).

Space for Personal Notes

Sub-Section [2.9.2]: Graph Effects of Temperature, Inert Gas, or Catalyst on Equilibrium Systems

Question 43 (1 mark)

A noble gas is added to an equilibrium system. Sketch the change to the system.



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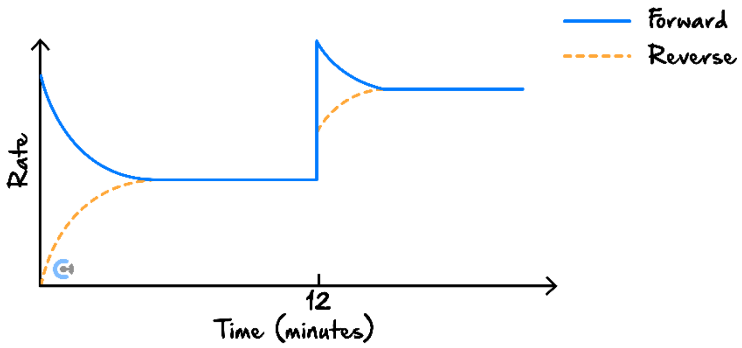
Sub-Section [2.9.3]: Find the Change made to the System from an Equilibrium Graph

Question 44 (1 mark)

A change is made to a gaseous equilibrium system at 12 minutes after the system is set up. The effect of the change is shown.



The change made to the system could be:

A. Reactants were added to the system.

B. The temperature was decreased.

C. A catalyst was added to the system.

D. The volume of the system was increased.

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Sub-Section [2.9.4]: Find Equilibrium Constant changes due to Temperature

Question 45 (1 mark)

Consider an endothermic equilibrium system for which the equation is unknown. The temperature of the system is increased. Which of the following would be observed once the system re-established equilibrium?

A. The rate of the reverse reaction only has decreased compared to before the temperature change.

B. The concentration of the reactants has increased.

C. The value of the equilibrium constant has increased.

D. The system feels colder.

Question 46 (7 marks)

The following system has been allowed to reach equilibrium.

$$4X(\text{aq}) \rightleftharpoons Y(\text{aq}) + Z(\text{aq}) \qquad \Delta < 0$$

At equilibrium, there are 4.00 mol of X and 6.00 mol of Y in 2 litre of solution. The value of the equilibrium constant is 0.1875 M^{-2}.

a. Calculate the concentration of Z. (2 marks)

$$\mathrm{K_c} = \frac{[Y][Z]}{[X]^4} \Longrightarrow [Z] = \frac{\mathrm{K_c}\,[X]^4}{[Y]} \ (1)$$

$$[Z] = \frac{0.1875\left(\frac{4}{2}\right)^4}{\frac{6}{2}} = 1.00\ M\ (2)$$

The chemist decides to change the temperature of the system and measures the concentration of Z to be 1.10 M after the change in temperature.

b. Explain whether the chemist increased or decreased the temperature of the system. (2 marks)

The concentration of Z increases and therefore the forward, exothermic reaction must be taking place (1). The exothermic reaction partially opposes a decrease in temperature according to Le Chatelier's principle and thus, the chemist decreases the temperature of the system (2).

c.

i. State the expression for the equilibrium constant. (1 mark)

$K_c = \frac{[Y][Z]}{[X]^4}\ M^{-2}$ (units not required)

ii. Calculate the new equilibrium constant. (2 marks)

R	$4X$	$\rightleftharpoons$	Y	+	Z
n(I)	4		6		2
n(C)	$-4x$ -0.8		$+x$ $+0.2$		$+x$ $+0.2$
n(E)	3.2		6.2		2.2
[E]	1.6M		3.1M		1.1M

at equilibrium,
[Z] = 1.10M
∴ n(E) for Z
= 1.10M × 2L
= 2.2 mol

$$K_c = \frac{[Y][Z]}{[X]^4} = \frac{(3.1)(1.1)}{(1.6)^4} = 0.520\ M^{-2}$$

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Sub-Section [2.9.5]: Find Optimum Operating Conditions in all Circumstances such as the Rate-Yield Conflict

Question 47 (1 mark)

A chemist is trying to maximise the yield of P in an industrial process involving the equilibrium system shown below.

$$2P(\text{aq}) \rightleftharpoons Q(\text{aq}) + 2R(\text{aq}) \quad \Delta H > 0$$

Which of the following could be applied to the system to maximise the yield of P?

A. Increasing the volume of the container.

B. Increasing the temperature of the system.

C. Reacting P with another species in a precipitate reaction.

D. Removing products as they are produced.

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Question 48 (8 marks)

A chemist is trying to increase the yield of Z in the following equilibrium system.

$$X(g) \rightleftharpoons 2Z(g) \qquad \Delta H > 0$$

a. Explain the effect of decreasing the temperature of the system on the value of the equilibrium constant. (2 marks)

A decrease in temperature is partially opposed by the exothermic reverse reaction. This results in an increase in $[X]$ and a decrease in $[Z]$ (1). Since the equilibrium constant is the ratio of products to reactants, the value of the equilibrium constant will decrease (2).

b. Compare the effect of decreasing the temperature of the system on the rate of the forward reaction at equilibrium and the rate of the reverse reaction at equilibrium. (2 marks)

Both the forward and reverse reaction, at dynamic equilibrium, must be occurring at the same rate (1). The effect of decreasing the temperature, therefore, is the same for rates of both the forward and reverse reactions (2). In that both rates will decrease.

c. Explain why increasing the volume of this system is not optimal for increasing the rate and yield of Z with explicit reference to collision theory and Le Chatelier's principle. (4 marks)

Increasing Volume will decrease the overall pressure of the system. According to Le Chatelier's principle, the system will partially oppose the change by increasing overall pressure, favoring the side with more particles (1). Thus, the forwards reaction which produces Z(g) is favored, increasing the yield of Z(g) (2). However, decreasing pressure results in less reacting particles being present in a given area. This will decrease the frequency of overall collisions, therefore decreasing the frequency of successful collisions in the correct orientation (3). Thus, the rate of reaction will decrease (4) and Z(g) will be produced at a slower rate. This is a rate-yield conflict and is thus not ideal.

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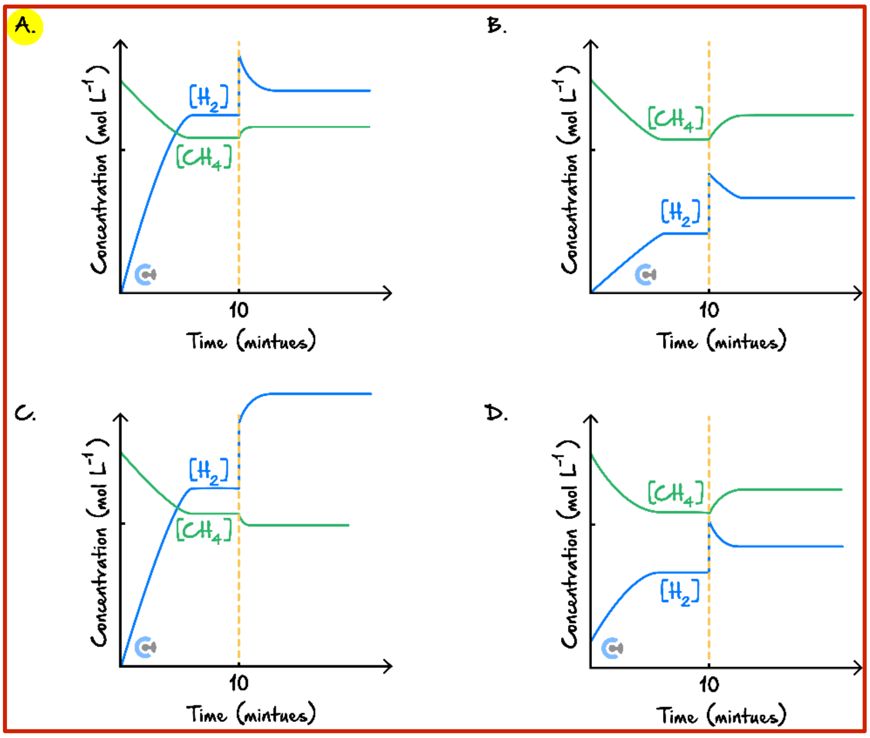
Section E: [2.6 - 2.9] - Overall (VCAA Qs) (57 Marks)

Question 49 (1 mark)

Inspired from VCAA CHEMISTRY Exam 2008
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2008chem2-w.pdf#page=3

Equal amounts of $CH_4(g)$ and $H_2O(g)$ are added to a reaction vessel and allowed to react. After 10 minutes, equilibrium has been reached. At that time, some H_2 is added to the mixture and equilibrium is re-established. Which one of the following graphs best represents the changes in the amounts of CH_4 and H_2 in the reaction mixture?



Question 50 (1 mark)

Inspired from VCAA CHEMISTRY Exam 2008
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2008chem2-w.pdf#page=3

The rate of a reaction generally increases with temperature.
The factor that has the **biggest effect** on the increase in reaction rate is that, with increasing temperature:

A. The activation energy of the reaction increases.

B. The activation energy of the reaction decreases.

C. The number of collisions between particles increases.

D. The proportion of particles with high kinetic energy increases.

Question 51 (1 mark)

Inspired from VCAA CHEMISTRY Exam 2023
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2023/2023chemistry-w.pdf#page=4

At 327°C, the equilibrium constant for the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ is $4.10\ M^{-2}$.
What is the equilibrium constant at 327°C for the reaction $2N_2(g) + 6H_2(g) \rightleftharpoons 4NH_3(g)$?

A. $8.20\ M^{-2}$

B. $8.20\ M^{-4}$

C. $16.8\ M^{-2}$

D. $16.8\ M^{-4}$

$N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$ $K = 4.10\ M^{-2}$

The original equation coefficients have been doubled, so the *K*value is raised to the power of 2 (squared).

$2N_2(g) + 6H_2(g) \leftrightharpoons 4NH_3(g)$ $K = (4.10\ M^{-2})^2$

$= 16.8\ M^{-4}$

Question 52 (1 mark)

Inspired from VCAA CHEMISTRY Exam 2022
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2022/NHT/2022ChemistryNHT-w.pdff#page=4

Hydrogen iodide, HI, can decompose to form hydrogen, H_2, and iodine, I_2. The equation for this reaction is given below.

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

Which one of the following statements about collision theory is correct for this reaction?

A. The frequency of collisions is independ

B. The frequency of collisions depends on

Reaction requires that reacting molecules collide in the right orientation and with energy greater than the activation energy.

C. The chance of a collision resulting in a reaction depends on the orientation of the molecules.

D. The chance of a collision resulting in a reaction depends only on the kinetic energy of the molecules.

Space for Personal Notes

Question 53 (1 mark)

Inspired from VCAA CHEMISTRY Exam 2021
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2021/NHT/2021chem-nht-w.pdff#page=4

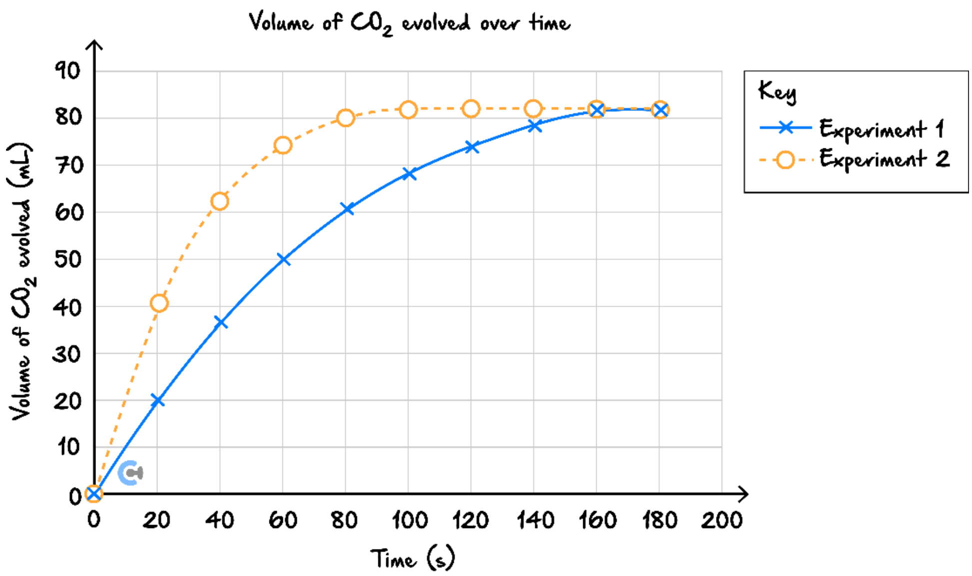
A student investigated the rate of reaction between dilute hydrochloric acid, HCl, and calcium carbonate, $CaCO_3$.
The student performed two experiments using two pieces of $CaCO_3$ of different shapes.

In each experiment, Experiment 1 and Experiment 2, 0.34 g of $CaCO_3$ and 12.0 mL of 0.50 M HCl was used and the volume of gas produced at SLC was measured at 10 s intervals.

The following reaction takes place:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

The molar mass of $CaCO_3$ is 100.1 $g\ mol^{-1}$.
The student plotted the results and produced the graph shown below.



At 160 s, the two lines meet.

This is because:

A. Both reactions have reached equilibrium.

B. The $CaCO_3$ has completely reacted and there is some acid remaining.

C. The acid has completely reacted and there is some $CaCO_3$ remaining.

D. Both reactions have stopped because all reactants have been consumed.

Space for Personal Notes

Question 54 (1 mark)

Inspired from VCAA CHEMISTRY Exam 2018
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2018/nht/2018chem-nht-w.pdf#page=3

Consider the following reaction.

$$O_2(g) + 2NO(g) \rightleftharpoons 2NO_2(g) \qquad K_c = 9.1 \times 10^6\ M^{-1} \text{ at } 450°C$$

Which one of the following statements are correct?

A. K_c is very large; the reaction is reversible.

B. Changing the pressure will affect the value of K_c.

C. K_c is very large; the reaction is effectively irreversible.

D. Changing the temperature will not affect the value of K_c.

Question 55 (1 mark)

Inspired from VCAA CHEMISTRY Exam 2023
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2023/NHT/2023chem-nht-w.pdf#page=3

Sulphur trioxide, SO_3, can be produced from oxygen, O_2, and sulphur dioxide, SO_2 in the following equilibrium reaction.

$$2SO_2(g) + O_2(g) \leftrightharpoons 2SO_3(g); \qquad \Delta H = -197\ kJ\ mol$$

Consider the following statements about the production of SO_3 in a closed system.

I. Increasing the temperature of the system will increase both the reaction rate and the equilibrium constant.

II. Increasing the volume of the system will reduce the reaction rate and reduce the equilibrium yield of SO_3.

III. Increasing the pressure of the system by adding nitrogen gas, N_2, will increase the equilibrium yield of SO_3.

IV. Adding a suitable catalyst to the system will have no effect on the equilibrium constant.

Which of the statements above is correct?

A. I and II only.

B. I and III only.

C. II and IV only.

D. III and IV only.

Question 56 (1 mark)

Inspired from VCAA CHEMISTRY Exam 2012
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2012/2012chem2-w.pdf#page=6

The following reaction is used in some industries to produce hydrogen.

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

Carbon monoxide, water vapour, carbon dioxide, and hydrogen were pumped into a sealed container that was maintained at a constant temperature of 200°C. After 30 seconds, the concentration of gases in the sealed container was found to be $[CO] = 0.1\ M, [H_2O] = 0.1\ M, [H_2] = 2.0\ M, [CO_2] = 2.0\ M$.
The equilibrium constant at 200°C for the above reaction is, K = 210.

Which one of the following statements about the relative rates of the forward reaction and the reverse reaction at 30 seconds is true?

A. The rate of the forward reaction is greater than the rate of the reverse reaction.

B. The rate of the forward reaction is equal to the rate of the reverse reaction.

C. The rate of the forward reaction is less than the rate of the reverse reaction.

D. There is insufficient information to allow a statement to be made about the relative rates of the forward and reverse reactions.

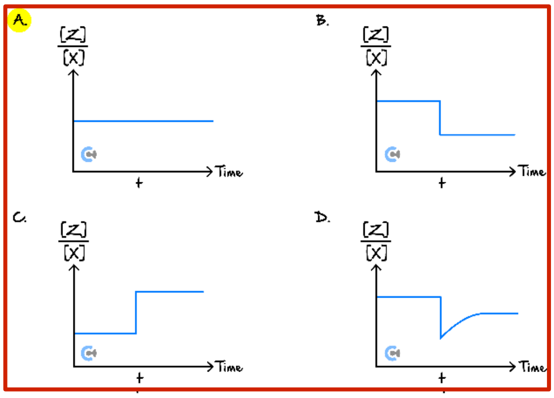
Question 57 (1 mark)

Inspired from VCAA CHEMISTRY Exam 2011
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2011chem2-w.pdf#page=2

Consider the following equilibrium.

$$2X(g) + Y(g) \rightleftharpoons 2Z(g) \qquad \Delta H \text{ is positive.}$$

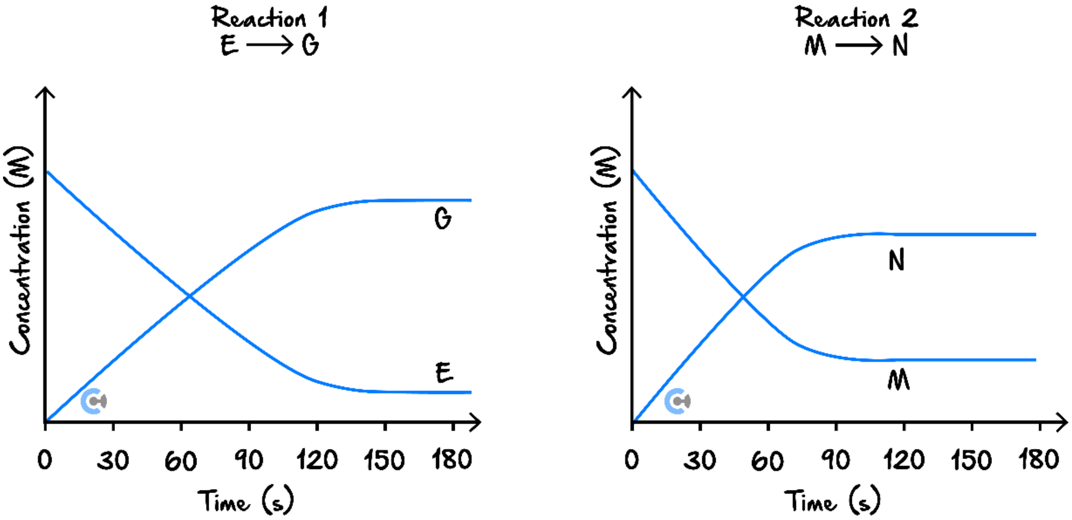
A catalyst is added to the equilibrium mixture at time, t. The temperature and volume of the mixture remain constant. Which one of the following graphs best represents the change in the equilibrium ratio $\frac{[Z]}{[X]}$?



Question 58 (1 mark)

Inspired from VCAA CHEMISTRY Exam 2019
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2019/NHT/2019chem-nht-w.pdf#page=8

The diagrams below show the concentration versus time graphs for two different reactions under identical conditions.
The graphs can be used to determine whether each reaction is reversible or irreversible and which reaction has a faster rate of reaction.



Which of the following conclusions matches the data presented in the graphs?

	Reaction 1	Rate of reaction
A.	Reversible.	Reaction 1 is slower than Reaction 2.
B.	Effectively irreversible.	Reaction 1 is faster than Reaction 2.
C.	Effectively irreversible.	Reaction 1 is slower than Reaction 2.
D.	Reversible.	Reaction 1 is faster than Reaction 2.

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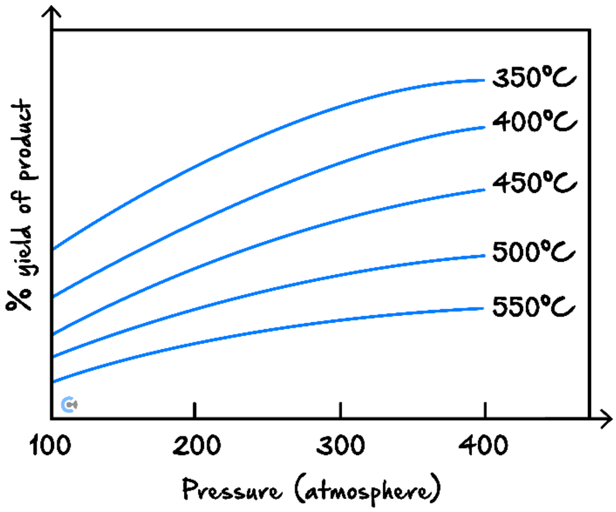
The following information applies to the two questions that follow.

Inspired from VCAA CHEMISTRY Exam 2010
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2010chem2-w.pdf#page=5

The graph below refers to the following gaseous reaction.

$$aA(g) + bB(g) \rightleftharpoons cC(g) + dD(g)$$

The effect of increasing pressure and temperature on the equilibrium yield of the products is shown in the graph below.



Question 59 (1 mark)

Which one of the following statements about the relative number of reactant and product molecules in the balanced equation is true?

A. The number of reactant molecules is equal to the number of product molecules.

B. The number of reactant molecules is greater than the number of product molecules.

C. The number of reactant molecules is less than the number of product molecules.

D. The relative number of reactant and product molecules cannot be determined from the data provided.

Question 60 (1 mark)

Which one of the following statements about this gaseous reaction is true?

A. The reaction is exothermic because the yield increases as the temperature increases.

B. The reaction is endothermic because the yield increases as the temperature increases.

C. The reaction is exothermic because the yield decreases as the temperature increases.

D. The reaction is endothermic because the yield decreases as the temperature increases.

The following information applies to the two questions that follow.

Inspired from VCAA CHEMISTRY Exam 2013
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2013/2013chem-w.pdf#page=8

$$2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g) \qquad \Delta H \text{ is positive.}$$

Question 61 (1 mark)

The equilibrium expression for this reaction is:

A. $\frac{2[NO][Cl_2]}{2[NOCl]}$

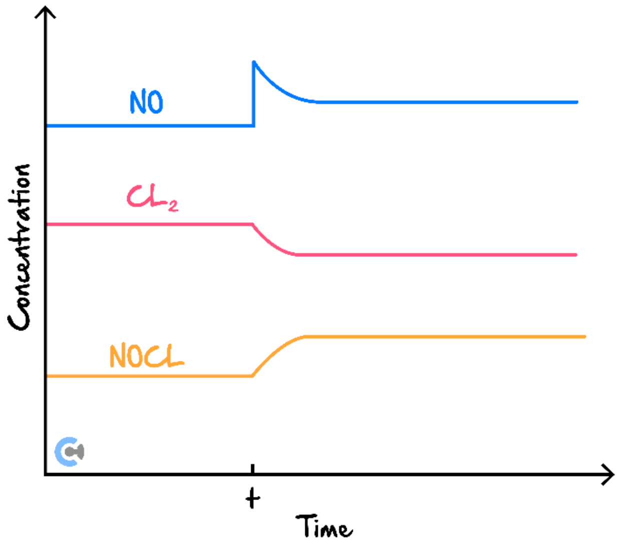
B. $\frac{[NO]^2[Cl_2]}{[NOCl]^2}$

C. $\frac{2[NOCl]}{2[NO][Cl_2]}$

D. $\frac{[NOCl]^2}{[NO]^2[Cl_2]}$

Question 62 (1 mark)

A concentration-time graph for this system is shown below.



What event occurred at time, t, to cause the change in equilibrium concentrations?

A. The pressure was decreased at a constant temperature.

B. The temperature was increased at a constant volume.

C. A catalyst was added at a constant temperature and volume.

D. Additional NO gas was added at a constant volume and temperature.

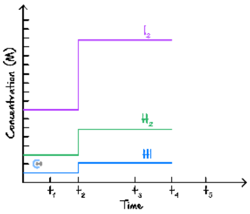
The following information applies to the two questions that follow.

Inspired from VCAA CHEMISTRY Exam 2021
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2021/2021chem-w.pdf#page=8

Hydrogen, H_2, and iodine, I_2, react to form hydrogen iodide, HI.

$$\frac{1}{2}H_2(g) + \frac{1}{2}I_2(g) \rightleftharpoons HI(g) \quad \Delta H = +25.9\ kJ\ mol^{-1}$$

The graph shows the concentrations H_2, I_2, and HI in a sealed container. One change was made to the equilibrium system at time, t_2.



Question 63 (1 mark)

Which one of the following statements is correct?

A. A catalyst was added at time t_2.

B. The amount of HI is greater at time t_3 compared with time t_1.

C. The rate of reaction producing HI is the same at time t_1 and time t_3.

D. The rate of production of HI at time t_3 is double the rate of production of H_2 at time t_3.

Question 64 (1 mark)

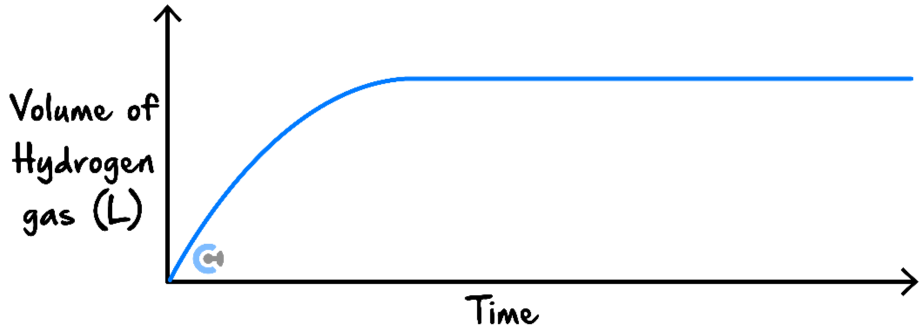
One change was made to the equilibrium system at time t_4, which altered the equilibrium constant. Equilibrium was re-established at time t_5. The rate of the reverse reaction at time t_5 was higher than at time t_3.
Which of the following options correctly shows the change in the equilibrium system from time t_3 to time t_5?

	Change from time t_3 to time t_5	
	Equilibrium Constant	**Total Chemical Energy**
A.	Increase	Increase
B.	Increase	Decrease
C.	Decrease	Increase
D.	Decrease	Decrease

Question 65 (4 marks)

Inspired from VCAA CHEMISTRY Exam 2008
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2008chem2-w.pdf#page=11

A 2.0 *g* piece of magnesium ribbon was added to a known volume of 2.0 *M* hydrochloric acid. The volume of hydrogen gas produced during the reaction was measured and recorded.
The graph below shows the result of this experiment.



a. Write an equation for the reaction between magnesium and hydrochloric acid. (2 marks)

Question 1a.

Marks	0	1	2	Average
%	18	29	54	1.4

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$* or

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

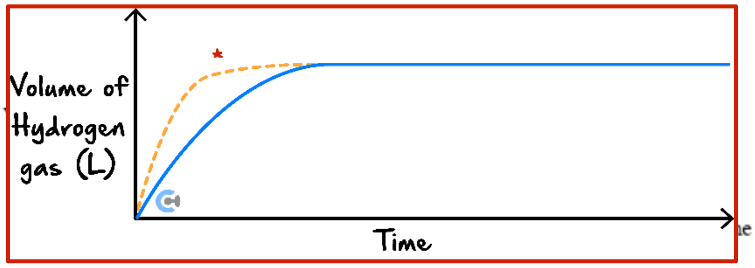
*One mark was awarded on this equation for 'all' states being correct.

b. In a second experiment, 2.0 *g* of magnesium **powder** was added to the same volume of 2.0 *M* hydrochloric acid as used in the first experiment.

On the axes above, sketch the expected graph of the volume of hydrogen against time for this second experiment. Give an explanation for the shape of your graph. (2 marks)

Question 1b.

Marks	0	1	2	Average
%	12	15	73	1.6



Powdered Mg has a **greater surface area** so the reaction occurs **faster***.

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Question 66 (7 marks)

Inspired from VCAA CHEMISTRY Exam 2021
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2021/2021chem-w.pdf

The reaction for the oxidation of sulphur dioxide, SO_2, is shown below.

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

a. 1.00 *mol* of SO_2 and 1.00 *mol* of oxygen, O_2, are placed into an evacuated, sealed 3.00 *L* container at 100°C. After the reaction reaches equilibrium, the container contains 20.0 *g* of sulfur trioxide, SO_3.

Calculate the equilibrium constant, K_c, for this reaction at 100°C. (4 marks)

Marks	0	1	2	3	4	Average
%	24	11	12	21	31	2.3

Multiple approaches were evident in students' responses to this question. The following is representative of one common approach:

n(SO_3) at equilibrium = 20.0 / 80.1 = 0.25 mol*

	n(SO_2) (mol)	n(O_2) (mol)	n(SO_3) (mol)
Initial	1.00	1.00	0.00
Change	-2x	-x	+2x
Final*	1.00-0.25 = 0.75	1.00-0.125 = 0.875	0.25

Therefore

$[SO_2]_{eqm} = {}^{0.75}/_{3.00} = 0.25$ M, $[O_2]_{eqm} = {}^{0.875}/_{3.00} = 0.292$ M, $[SO_3]_{eqm} = {}^{0.25}/_{3.00} = 0.083$ M*

$K = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{0.083^2}{0.25^2 \times 0.292} = 0.38\ M^{-1}$*

In order to gain full marks, students were required to submit the final calculated value with the correct units. Consequential marks were awarded when students were able to show clear calculations.

b. The volume of the closed container is doubled.

Describe the effect that this has on the concentration of SO_2 from the time just before the volume was changed until after the system re-established its equilibrium. (3 marks)

Marks	0	1	2	3	Average
%	35	30	19	16	1.2

Responses had to be linked to the effect on the concentration of SO_2, as follows:

- As the volume doubles, the concentration of SO_2 halves.
- As the system moves to re-establish equilibrium, the concentration of SO_2 gradually increases.
- Once equilibrium is reached, the final concentration of SO_2 is still lower than the original concentration (before volume increase).

Although it was not explicitly required, some students included a concentration-time graph that assisted them to communicate their understanding.

The most common issue for students was a lack of clarity when explaining the effect. A lot of students discussed the cause that underpinned this change, which was not required by the question. This highlights the need for students to make sure that they clearly read the stem of each question and apply a response that is appropriate to what is required of them.

Question 67 (8 marks)

Inspired from VCAA CHEMISTRY Exam 2018
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2018/2018chem-w.pdf

A student is investigating the following reaction system.

$$\underset{\text{brown}}{2NO_2(g)} \rightleftharpoons \underset{\text{colourless}}{N_2O_4(g)} \quad \Delta H < 0$$

a. The reaction system can be observed in a sealed test tube, which allows the student to investigate the impact of temperature on the equilibrium position of the reaction.

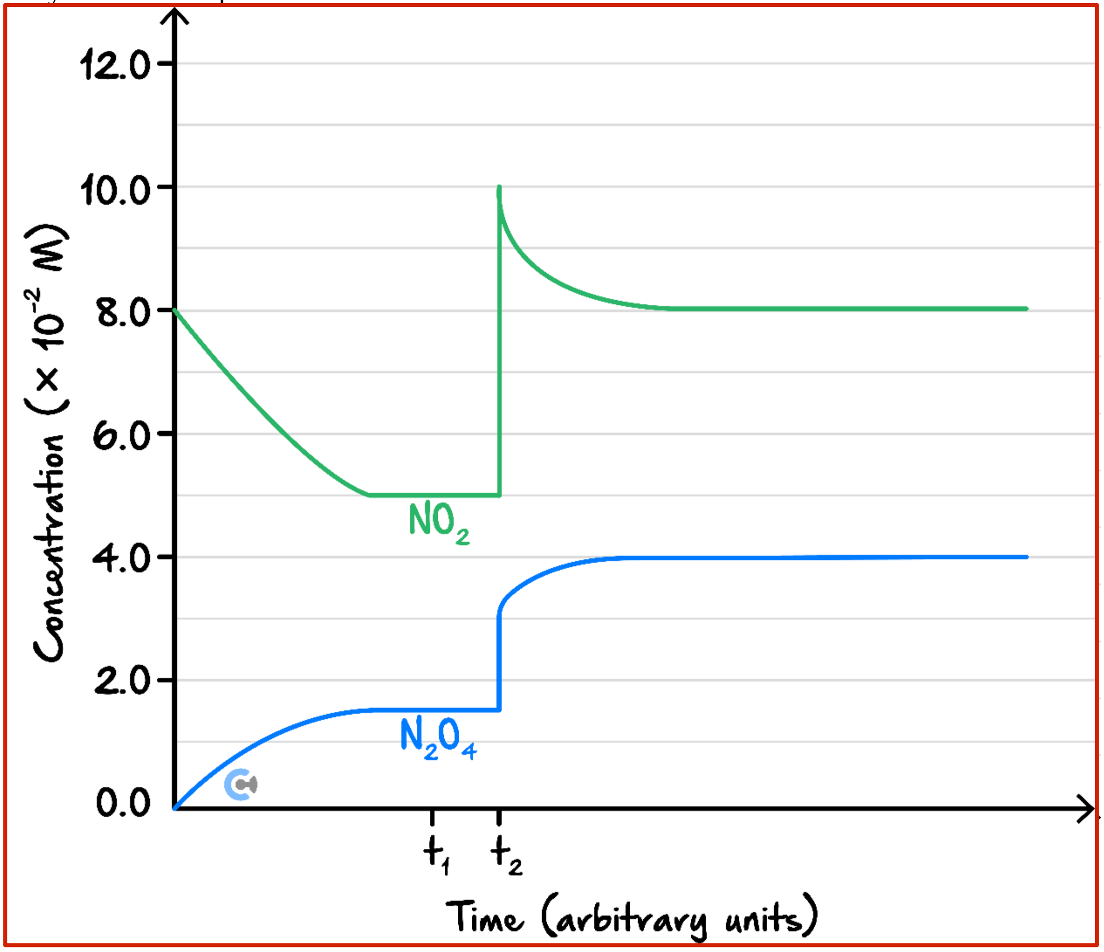
State the colour change expected when the student places the sealed test tube of the gas mixture in a beaker of hot water. Explain why this colour change occurs. (3 marks)

Question 7a.

Marks	0	1	2	3	Average
%	26	17	28	29	1.6

- The gas mixture goes darker brown.
- The forward reaction is exothermic ($\Delta H < 0$), hence the reverse reaction is favoured at higher temperatures.
- As the reverse reaction is favoured, $[NO_2]$ increases and so the equilibrium mixture is darker at higher temperature.

b. Below is the concentration versus time graph for the reaction system. The graph was produced using secondary data at a temperature of 22°C.



i. Time t_1, is shown in the graph above.

Calculate the equilibrium constant at time t_1. (2 marks)

Question 7bi.

Marks	0	1	2	Average
%	22	48	30	1.1

$K = [N_2O_4] / [NO_2]^2$

$= 1.5 \times 10^{-2}/(5.0 \times 10^{-2})^2$

$= 6.0\ M^{-1}$

Most students identified the equilibrium law but many either misread the concentrations from the graph or did not include the appropriate unit for the value of the equilibrium constant. The ability to read graphical data accurately is a fundamental skill.

ii. At time t_2, the volume of the system was halved, keeping the temperature at 22°C.

Continue the graph to show how this change would affect the reaction system and how the system would respond to this change until equilibrium is restored. (3 marks)

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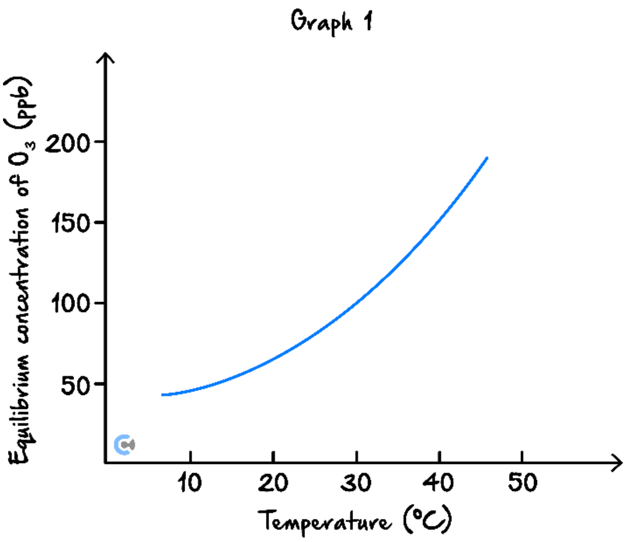
Question 68 (8 marks)

Inspired from VCAA CHEMISTRY Exam 2022
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2022/NHT/2022ChemistryNHT-w.pdf#page=20

In the upper atmosphere, ozone, $O_3(g)$, is formed from oxygen, $O_2(g)$, in the presence of ultraviolet (UV) light. An equation that represents this chemical reaction is given below.

$$3O_2(g) \underset{\text{light}}{\overset{\text{UV}}{\rightleftharpoons}} 2O_3(g)$$

Graph 1 shows the effect of temperature on the equilibrium concentration of $O_3(g)$ in a sealed container containing only $O_2(g)$ and $O_3(g)$. The container is clear and exposed to UV light.



a.

i. With reference to Graph 1, state whether the forward reaction is endothermic or exothermic. (1 mark)

Since the $[O_3]$ (product) increases with temperature, the forward reaction is endothermic.

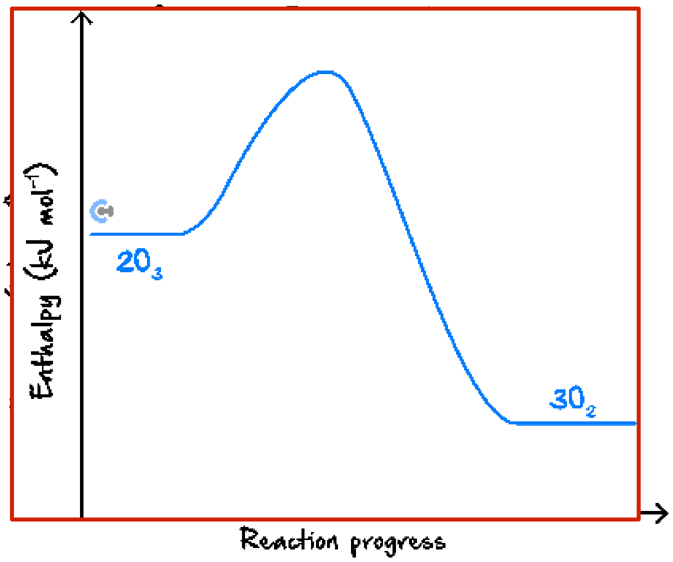
ii. Explain your answer to **part a.i.** (2 marks)

Students need to incorporate Le Chatelier's principle in their explanation.

As the temperature increases, there is a net shift to the right to produce more O_3 / forward reaction is favoured. *

System moves to partially oppose the increase in temperature by favouring the energy-absorbing (endothermic) reaction. *

iii. Decomposition of O_3 to produce O_2 is the reverse reaction.
Sketch the energy profile for the decomposition of O_3 on the axes provided in Graph 2. (1 mark)



b. 7.50×10^{-2} mol of O_2 was placed in an evacuated and sealed 3.00 L container at 30°C. This clear container was exposed to UV light. At equilibrium, 1.56×10^{-7} mol of O_3 had formed in the container.

Calculate the equilibrium constant at 30°C. (4 marks)

	$3O_2$	$\leftrightharpoons$	$2O_3$
	$n(O_2)$ mol		$n(O_3)$ mol
Initially	7.50×10^{-2}		0
Change	$-x$		$+\frac{2x}{3}$
	-2.34×10^{-7}		$+ 1.56 \times 10^{-7}$ *
Equilibrium	0.0750		1.56×10^{-7}

$[O_2]_{eqm} = {}^{0.0750}/_{3.00} = 0.0250$ M,

$[O_3]_{eqm} = 1.56 \times 10^{-3} / 3.00 = 5.20 \times 10^{-8}$ M *

$K = [O_3]^2 / [O_2]^3$ * $= (5.20 \times 10^{-8})^2 / 0.0250^3$

$= 2.70 \times 10^{-15} / 1.56 \times 10^{-5}$

$= 1.73 \times 10^{-10}$ M^{-1} *

To gain full marks, students were required to submit the final calculated value with the correct units:

- 1 mark for change in moles of O_2
- 1 mark for equilibrium concentrations of O_2 and O_3
- 1 mark for equilibrium expression
- 1 mark for substitution of concentrations into equilibrium expression and final answer.

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Question 69 (8 marks)

Inspired from VCAA CHEMISTRY Exam 2022
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2022/NHT/2022ChemistryNHT-w.pdf#page=20

Nitrosyl chloride, NOCl, is a highly toxic gas used in the chemical industry as an oxidising agent.
The formation reaction of NOCl from nitrogen monoxide, NO, and chlorine, Cl_2, is:

$$2\text{NO}(g) + \text{Cl}_2(g) \rightleftharpoons 2\text{NOCl}(g) \quad \Delta H > 0$$

This reaction forms an equilibrium above 100°C.
A scientist conducted two experiments on the equilibrium reaction of NOCl. The initial experiments were conducted in evacuated and sealed 4 L containers at 150°C.

Experiment 1: 2 mol of NOCl was injected into a previously evacuated, sealed 4 L container.

Experiment 2: 4 mol of NOCl was injected into another previously evacuated, sealed 4 L container.

a.

i. Which experiment had the highest initial rate of production of Cl_2? Circle the correct response below. Justify your answer. (2 marks)

Experiment 1 | Rates are equal | Experiment 2

Experiment 2 has a higher initial concentration of NOCl molecules, so there will be a higher frequency of collisions (higher number of collisions per unit time) and so a higher frequency of successful collisions (higher number of successful collisions per unit time) producing a faster rate of reaction/rate of production of Cl_2.

One mark each was awarded for:

- Experiment 2
- correctly justifying the higher initial rate in Experiment 2.

ii. If, for Experiment 1, the concentrations of NOCl and NO were equal at equilibrium, [NOCl] = [NO], then what conclusion could be made about the relative concentrations of NOCl and NO in Experiment 2 at equilibrium? Justify your answer. (2 marks)

[NOCl] is higher than [NO].

Experiment 2 is like adding an extra 2 mol NOCl to the equilibrium in Experiment 1. This immediately increases the ratio $[NOCl]^2/\{[NO]^2[Cl_2]\}$. Because the reaction shifts to the left the $[Cl_2]$ will be higher at equilibrium in Experiment 2. However, since $[NOCl]^2/\{[NO]^2[Cl_2]$ at equilibrium must be the same at both equilibria, the ratio $[NOCl]^2/[NO]^2$ must increase and so [NOCl] must be greater than [NO] at equilibrium in Experiment 2.

$K = [NOCl]^2/([NO][Cl_2])$

Experiment 1 [NOCl] = [NO] so $K = 1/[Cl_2]$

In experiment 2, the higher initial [NOCl] leads to a higher $[Cl_2]$ at equilibrium.

Since $[NOCl]^2/([NO][Cl_2])$ is constant at equilibrium the ratio $[NOCl]^2 / [NO]^2$ must be larger in Experiment 2, so [NOCl] is greater than [NO]

Alternatively $K = [NO]^2[Cl_2]/[NOCl]^2$

Experiment 1 [NOCl] = [NO] so $K = [Cl_2]$

Since $[NO]^2[Cl_2]/[NOCl]^2$ is constant the ratio

$[NO]^2/[NOCl]^2$ must be smaller so [NOCl] is greater than [NO].

One mark each was awarded for:

- [NOCl] > [NO]
- for a valid explanation.

b. 2 mol of an inert gas is injected into the container in Experiment 1. The temperature is kept at 150°C. What effect will this have on the rate of production of the Cl_2 in the container? Circle the correct response below. Justify your answer using collision theory. (2 marks)

Decrease No change Increase

No change

There are more collisions due to added particles but not more successful collisions since collisions involving inert gas particles do not affect equilibrium. Inert gas molecules are not part of the equilibrium reaction.

c. The temperature for Experiment 2 is increased to 200°C. Explain the effect on the equilibrium concentration of NOCl in the reaction. (2 marks)

Endothermic reaction so an increase in temperature shifts equilibrium to the right

Hence [NOCl] will increase.

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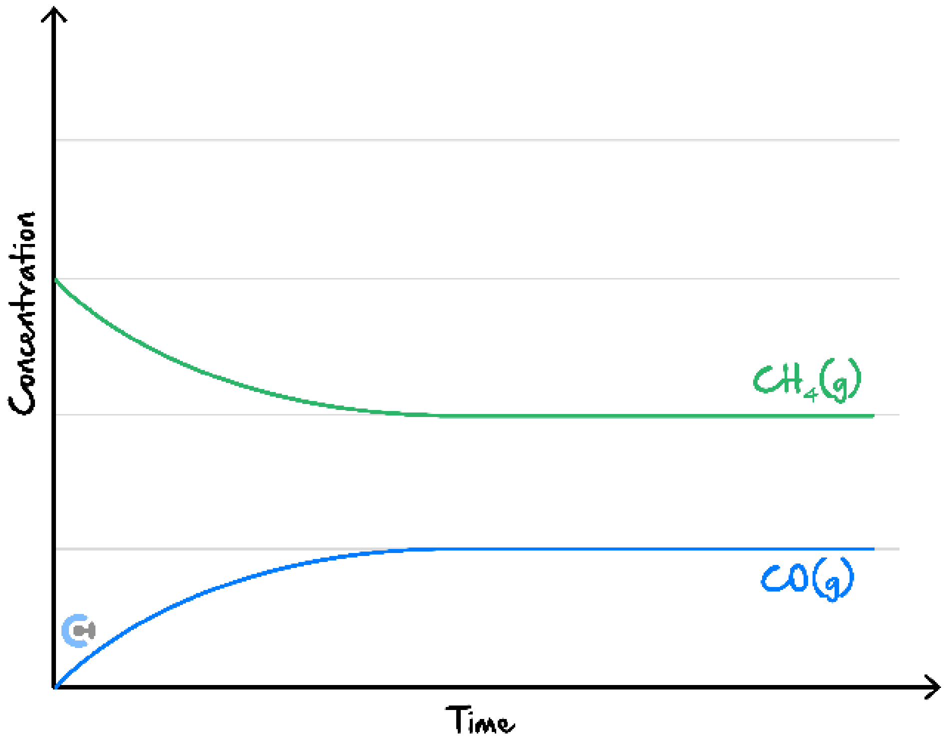
Question 70 (6 marks)

Inspired from VCAA CHEMISTRY Exam 2007
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/chem1_exam_07.pdf#page=10

Carbon monoxide and hydrogen can be produced from the reaction of methane with steam according to the equation:

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

Some methane and steam are placed in a closed container and allowed to react at a fixed temperature. The following graph shows the change in concentration of methane and carbon monoxide as the reaction progresses.



a.

i. On the graph above, draw a line to show the change in concentration of hydrogen gas as the reaction progresses. **Label this line**. (1 mark)

ii. On the graph above, draw a line to show how the formation of carbon monoxide would differ over time in the presence of a catalyst. **Label this line**. (1 mark)

b. The rates of chemical reactions may be explained using the collision theory model. Indicate whether the following statements about rates and the collision theory model are true or false by placing ticks in the appropriate boxes. (3 marks)

Statement	True	False
i. Endothermic reactions are always slower than exothermic reactions.		
ii.		
iii.		
iv.		
v. Increasing the temperature increases the fraction of collisions with energy above the activation energy.		

Marks	0	1	2	3	Average
%	1	16	42	41	2.3

Statement	True	False
i. Endothermic reactions are always slower than exothermic reactions.		✓
ii. All particles have the same kinetic energy at a fixed temperature.		✓
iii. Reactant particles need to collide with sufficient energy at a fixed temperature.	✓	
iv. The rate of a reaction at a constant temperature increases as the reaction proceeds.		✓
v. Increasing the temperature increases the fraction of collisions with energy above the activation energy.	✓	

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