1. This question is about covalent compounds of nitrogen.

Ammonia, NH_3, is manufactured by reacting nitrogen and hydrogen gases. This is a reversible reaction and the equilibrium is shown below.

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

i. This is an example of a dynamic equilibrium.

State **2** features of a dynamic equilibrium.

1 _____

_____

2 _____

_____ **[2]**

ii. State and explain the conditions of temperature and pressure that would produce a large equilibrium yield of NH_3.

_____

_____

_____

_____

_____

_____

_____ **[3]**

2. The reversible reaction below is at equilibrium.

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

What is the expression for K_c?

A $\frac{[SO_2(g)]^2\,[O_2(g)]}{[SO_3(g)]^2}$

B $\frac{[SO_3(g)]^2}{[SO_2(g)]^2\,[O_2(g)]}$

C $\frac{2[SO_2(g)] + [O_2(g)]}{2[SO_3(g)]}$

D $\frac{2[SO_3(g)]}{2[SO_2(g)] + [O_2(g)]}$

Your answer ☐ **[1]**

3. This question is about two oxides of sulfur: sulfur dioxide, SO_2, and sulfur trioxide, SO_3.

SO_3 decomposes to form SO_2 and O_2, as shown in **Equilibrium 18.1**.

$$SO_3(g) \rightleftharpoons SO_2(g) + \tfrac{1}{2}O_2(g) \qquad \Delta H = +99 \text{ kJ mol}^{-1} \qquad \textbf{Equilibrium 18.1}$$

i. 2.25 moles of SO_3 is heated to 550 °C in the presence of a catalyst and the resulting mixture allowed to reach equilibrium.

The equilibrium mixture contains 0.900 mol of SO_2 and the total pressure is 2.80 atm.

Calculate the numerical value for K_p for **Equilibrium 18.1** under these conditions and state the units of K_p.

Give your answer to **3** significant figures.

K_p = ...

units ..... **[5]**

ii. The numerical values of K_p for **Equilibrium 18.1** at temperatures T_1 and T_2 are shown below.

Temperature	K_p
T_1	3.3×10^{-5}
T_2	7.7×10^{-2}

Explain why T_2 is a higher temperature than T_1.

[2]

iii. Suggest how the value of K_p would change if the reaction was repeated with no catalyst added and the pressure of the system increased.

Tick (✓) one box in each row.

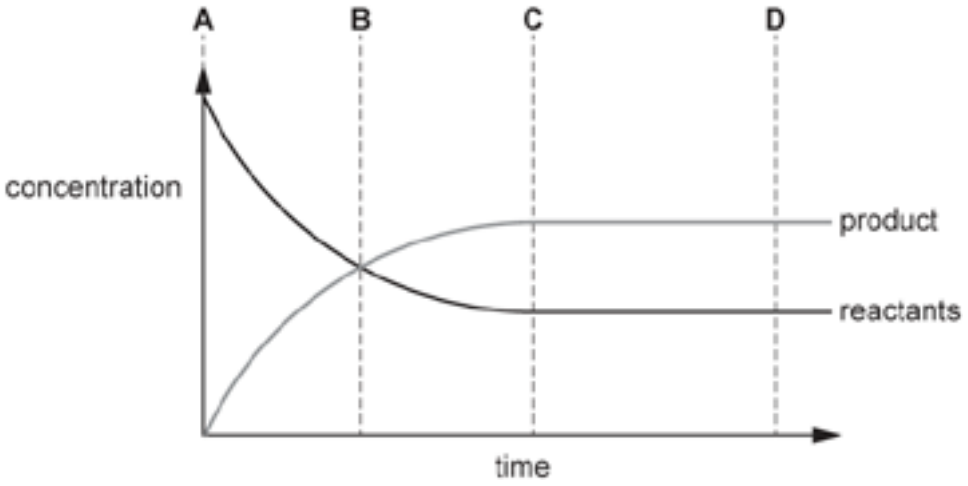
Change	Decrease	No change	Increase
No catalyst			
Increased pressure			

[2]

4. The reversible reaction between hydrogen and iodine to form hydrogen iodide is $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

The graph shows how the concentrations of the reactants and product change as the reaction reaches a dynamic equilibrium.

At which point on the graph is the equilibrium reached?



Your answer ☐ [1]

5(a). This question is about oxides of nitrogen.

An investigation is carried out on the equilibrium system shown below.

$2NO_2(g) \rightleftharpoons N_2O_4(g)$ $\Delta H = -57.4$ kJ mol^{-1}

i. A sealed flask containing 6.00 moles of $NO_2(g)$ is heated to a constant temperature and allowed to reach equilibrium.

The equilibrium mixture contains 5.40 mol of $NO_2(g)$, and the total pressure is 5.00 atm.

Determine the value of K_p and give your answer to **3** significant figures.

Include an expression for K_p and the units of K_p in your answer.

K_p = ... units .....**[5]**

ii. The sealed flask in **(a)(i)** is then heated to a higher temperature at an increased pressure. The system is allowed to reach equilibrium again.

Explain why it is difficult to predict how these changes in reaction conditions affect the amount of $N_2O_4(g)$ formed at equilibrium.

[3]

(b). N_2O_4 reacts fully with oxygen to form a different oxide of nitrogen, oxide **A**, as the only product.

Oxide **A** is collected and cooled to 75.0 °C at a pressure of 101 kPa.

Under these conditions, oxide **A** is a gas that occupies a volume of 74.0 cm^3 and has a mass of 0.280 g.

Calculate the molar mass of oxide **A** and suggest its molecular formula.

molar mass = ... g mol^{-1}
molecular formula = ...
[5]

6. Chlorine trifluoride can be decomposed into its elements forming the equilibrium mixture below.

$$\underset{\text{colourless gas}}{2ClF_3(g)} \rightleftharpoons \underset{\text{green gas}}{Cl_2(g)} + \underset{\text{yellow gas}}{3F_2(g)} \qquad \Delta H = +318 \text{ kJ mol}^{-1}$$

Which statement(s) is/are correct?

1 The decomposition is a redox reaction.
2 When the equilibrium mixture is cooled, the colour fades.
3 The decomposition has a negative entropy change.

A 1, 2 and 3
B Only 1 and 2
C Only 2 and 3
D Only 1

Your answer ☐ **[1]**

7. Bromine, Br_2, can be produced by the reaction:

$5Br^-(aq) + BrO_3^-(aq) + 6H^+(aq) \rightarrow 3Br_2(aq) + 3H_2O(l)$

A student investigates the rate of this reaction by carrying out four experiments at the same temperature. The student's results are shown below.

Experiment	$[Br^-]$ / mol dm^{-3}	$[BrO_3^-]$ / mol dm^{-3}	[H+] / mol dm^{-3}	Initial rate / mol dm^{-3} s^{-1}
1	2.00×10^{-2}	1.20×10^{-1}	8.00×10^{-2}	2.52×10^{-4}
2	6.00×10^{-2}	1.20×10^{-1}	8.00×10^{-2}	7.56×10^{-4}
3	4.00×10^{-2}	6.00×10^{-2}	8.00×10^{-2}	2.52×10^{-4}
4	2.00×10^{-2}	6.00×10^{-2}	4.00×10^{-1}	3.15×10^{-3}

Explain how the reaction orders can be determined from the student's results, and determine the rate equation and rate constant for this reaction.

[6]

8(a). The reaction between sulfur dioxide, $SO_2(g)$ and oxygen, $O_2(g)$, to form sulfur trioxide, $SO_3(g)$, is a key step in the industrial manufacture of sulfuric acid.

This is a reversible reaction, shown in **Equilibrium 24.1**:

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ $\Delta H = -197$ kJ mol^{-1} **Equilibrium 24.1**

Why is **Equilibrium 24.1** a homogeneous equilibrium?

[1]

(b). Le Chatelier's principle can be used to predict how different conditions affect the equilibrium position in **Equilibrium 24.1**.

Explain how changing pressure, temperature and using a catalyst affect the equilibrium yield of SO_3.

In your answer, use le Chatelier's principle and other chemical concepts, where appropriate.

[5]

(c). A mixture of $SO_2(g)$ and $O_2(g)$ is allowed to reach equilibrium at a constant temperature.

The equilibrium concentrations are shown in the table.

Substance	Equilibrium concentration / mol dm $^{-3}$
$SO_2(g)$	3.0×10^{-3}
$O_2(g)$	3.5×10^{-3}
$SO_3(g)$	5.0×10^{-2}

i. Write the expression for K_c and calculate the numerical value for K_c in **Equilibrium 24.1** at this constant temperature.

Give your answer to an **appropriate** number of significant figures and in **standard form**.

K_c = ... $dm^3\ mol^{-1}$ **[2]**

ii. In the industrial production of SO_3, an excess of $O_2(g)$ is used rather than a 2:1 proportion of $SO_2(g)$ to $O_2(g)$ which would match the stochiometry in **Equilibrium 24.1**.

Suggest, in terms of equilibrium, why an excess of $O_2(g)$ is used industrially.

[1]

9. Which prediction can be made using le Chatelier's principle?

A The effect of a catalyst on the reaction rate.
B The effect of a catalyst on the equilibrium position.
C The effect of temperature on the reaction rate.
D The effect of concentration on the equilibrium position.

Your answer ☐ **[1]**

10. Four equilibrium reactions are set up

The concentration of each gas in the equilibrium mixtures is 0.1 mol dm^{-3}.

Which equilibrium has a numerical K_c value of 0.01?

A $CH_4(g) + 2H_2O(g) \leftrightharpoons CO_2(g) + 4H_2(g)$
B $N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$
C $H_2(g) + I_2(g) \leftrightharpoons 2HI(g)$
D $2NO_2(g) \leftrightharpoons N_2O_4(g)$

Your answer ☐ **[1]**

11(a). This question is about equilibria involving hydrogen.

***** Hydrogen is used industrially to manufacture ammonia.

The equilibrium is shown below.

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -92 kJ\ mol^{-1}$ **Equilibrium 20.1**

1.20 mol $N_2(g)$ is mixed with 3.60 mol $H_2(g)$ in a 8.00 dm^3 container.

The mixture is heated to 550 °C with an iron catalyst and allowed to reach equilibrium.

The equilibrium mixture contains 0.160 mol of NH_3.

Determine the equilibrium constant K_c for **Equilibrium 20.1**, and explain why the operational conditions used by industry may be different from those required for a maximum equilibrium yield of ammonia.

[6]

(b). In industry, hydrogen is used to reduce the iron oxide Fe_3O_4 as shown in **Equilibrium 20.2**.

The reaction is carried out at 500 °C.

$Fe_3O_4(s) + 4H_2(g) \rightleftharpoons 3Fe(s) + 4H_2O(g)$ **Equilibrium 20.2**

i. When the temperature is decreased, the value of K_p decreases.
Determine whether the forward reaction is exothermic or endothermic.
Explain your answer.

[1]

ii. Two students are discussing the effect of pressure on the equilibrium position of **Equilibrium 20.2**.

Student 1 says:

"There are more moles of products than reactants, so increasing the pressure will shift the equilibrium to the left hand side."

Student 2 disagrees.

Determine which student is correct. Justify your answer.

[1]

12. This question is about the manufacture of hydrogen, H_2.

In industry, hydrogen is manufactured from methane, as shown in **Equilibrium 4.1**.

$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$ $\Delta H = +206$ kJ mol^{-1} **Equilibrium 4.1**

The industrial process is carried out at 15 atmospheres pressure and at a temperature of 800°C using an excess of steam. A nickel catalyst is used.

i. ***** Explain why these conditions are used industrially.

[6]

ii. A chemist mixes $CH_4(g)$ and $H_2O(g)$ and leaves the mixture to reach equilibrium.

$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$ $\Delta H = +206$ kJ mol^{-1} **Equilibrium 4.1**

The equilibrium mixture contains the following concentrations.

Substance	Concentration/mol dm^{-3}
$CH_4(g)$	0.111
$H_2O(g)$	0.682
$CO(g)$	0.510
$H_2(g)$	1.530

Write an expression for the equilibrium constant, K_c, for **Equilibrium 4.1** and calculate the numerical value of K_c.

Give your answer to **3** significant figures.

K_c = ... mol^2 dm^{-6} **[2]**

13(a). Two students plan to investigate **Equilibrium 4.1**, shown below.

$CoCl_4^{2-}(aq) + 6H_2O(l) \rightleftharpoons [Co(H_2O)_6]^{2+}(aq) + 4Cl^-(aq)$ **Equilibrium 4.1**

blue pink

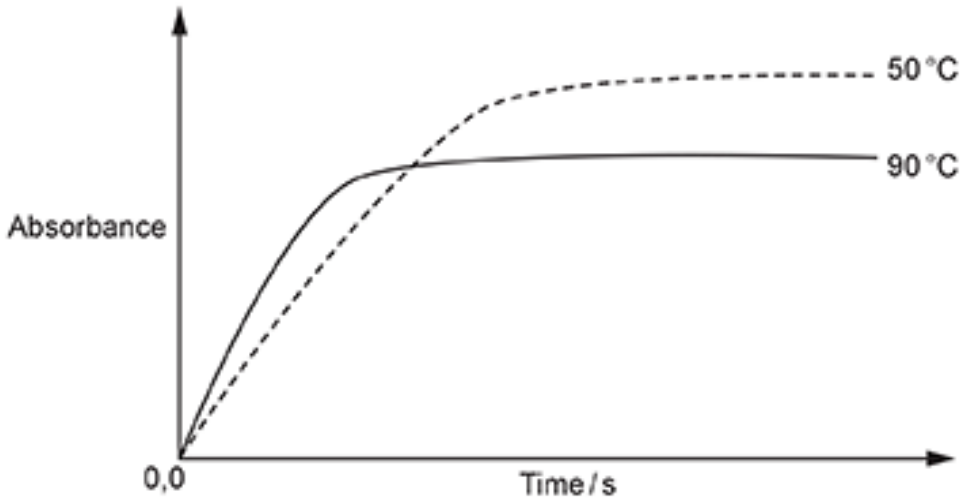
The students are supplied with the equilibrium mixture in **Equilibrium 4.1** at room temperature.

- One student heats 20 cm^3 of the mixture to 50°C.
- The other student heats 20 cm^3 of the mixture to 90°C.

The students use colorimetry to observe how the colour of the equilibrium mixture changes over time.

- The colorimeter is set up so that the greater the absorbance, the greater the concentration of $[Co(H_2O)_6]^{2+}$.
- The initial absorbance is set to zero.
- The absorbance is recorded every 30 seconds.

The students plot the graph below from the results of the experiment.



Use the graph and relevant chemical theory to answer the following. Include all reasoning:

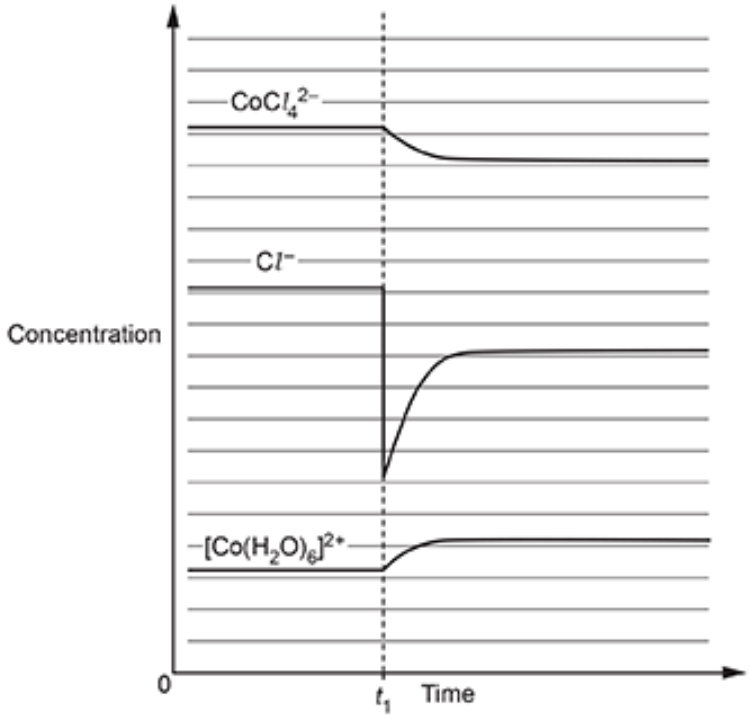
- Explain the different initial rates at 50°C and 90°C.
- Predict the sign of ΔH for the forward reaction in **Equilibrium 4.1**.

[4]

(b). The students investigate how addition of aqueous silver nitrate, $AgNO_3(aq)$, affects the equilibrium position in **Equilibrium 4.1**.

The graph shows the changes in the equilibrium concentrations of $CoCl_4^{2-}$, Cl^- and $[Co(H_2O)_6]^{2+}$ after addition of the $AgNO_3(aq)$.

The $AgNO_3(aq)$ is added at time = t_1



i. Explain why the Cl^- concentration drops sharply at time = t_1.

[1]

ii. Explain the changes in concentration of $CoCl_4^{2-}$, Cl^- and $[Co(H_2O)_6]^{2+}$ after time = t_1.
Refer to **Equilibrium 4.1** in your answer.

[3]

END OF QUESTION PAPER