1. The following reaction is used in industry to make sulfur trioxide gas, SO_3.

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

This preparation is carried out in the presence of a catalyst.

Explain the conditions of temperature and pressure that could be used to obtain the maximum equilibrium yield of sulfur trioxide.

Discuss the importance of a compromise between equilibrium yield and reaction rate when deciding the operational conditions for this process.

[6]

2. The reaction of ammonia, NH_3, with oxygen to form nitrogen monoxide, NO, is an important industrial process.

The equation for this reaction is shown in **equilibrium 4.1** below.

$$4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g) \qquad \Delta H = -905 \text{ kJ mol}^{-1} \qquad \textbf{Equilibrium 4.1}$$

Predict the conditions of temperature and pressure for a maximum equilibrium yield of nitrogen monoxide in **equilibrium 4.1**.

- Explain your prediction in terms of le Chatelier's principle.
- State and explain how these conditions could be changed to achieve a compromise between equilibrium yield, rate and other operational factors.

[5]

3. Ammonia, NH_3, is manufactured by the chemical industry from nitrogen and hydrogen gases.

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

- An iron catalyst is used which provides several benefits for sustainability.
- The chemical industry uses operational conditions that are different from the conditions predicted to give a maximum equilibrium yield.

* Use your understanding of Chemistry to explain the above statements.

Your response should be well-developed, showing a line of reasoning which is clear and logically structured.

[6]

4. In the manufacture of methanol, carbon dioxide and hydrogen are reacted together in the reversible reaction shown below.

$$CO_2(g) + 3H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g) \qquad \Delta H = -49 \text{ kJ mol}^{-1}$$

High pressures and low temperatures would give a maximum equilibrium yield of methanol.

i. Explain this statement in terms of le Chatelier's principle.

[3]

ii. Explain why the actual conditions used by the chemical industry might be different.

[2]

6. Hydrogen gas is manufactured by the chemical industry using the reaction of methane and steam. This is a reversible reaction, shown in **equilibrium 20.1** below.

equilibrium 20.1 $CH_4(g) + H_2O(g) \rightleftharpoons 3H_2(g) + CO(g) \qquad \Delta H = +210 \text{ kJ mol}^{-1}$

Explain, in terms of le Chatelier's principle, the conditions of pressure and temperature for a maximum yield of hydrogen from **equilibrium 20.1**, and explain why the operational conditions used by the chemical industry may be different.

[4]