1. The following series of steps describes a reaction mechanism for a chemical reaction:

Step 1. \[ \text{H}_2\text{O}_2 + \text{H}^+ \rightarrow \text{H}_3\text{O}^+ \] fast

Step 2. \[ \text{H}_3\text{O}^+ + \text{I}^- \rightarrow \text{H}_2\text{O} + \text{HOI} \] slow

Step 3. \[ \text{HOI} + \text{I}^- \rightarrow \text{OH}^- + \text{I}_2 \] fast

Step 4. \[ \text{OH}^- + \text{H}^+ \rightarrow \text{H}_2\text{O} \] fast

Step 5. \[ \text{I}_2 + \text{I}^- \rightarrow 2\text{I}^- \] fast

Write the equation for the overall reaction and identify all the reaction intermediates. Increasing the concentration of which reactant will greatly increase the rate of the reaction? Explain.

\[ \text{H}_2\text{O}_2 + 2\text{H}^+ + 3\text{I}^- \rightarrow \frac{1}{2}\text{I}_2 + 2\text{H}_2\text{O} + 3\text{I}^- \]

Intermediates are: \( \text{H}_3\text{O}^+ \), \( \text{HOI} \), \( \text{OH}^- \), and \( \text{I}_2 \). Increasing concentration of \( \text{I}^- \) will increase rate as it is a

2. Consider the following potential energy diagram:

(a) On the diagram, label the change in enthalpy and the activation energy for the reverse reaction.

(b) Give the values for:
   i) the energy of the activated complex. \( 40 \) kJ
   ii) \( \Delta H \) for the forward reaction. \( -15 \) kJ

3. Describe TWO ways, other than the use of a catalyst, to increase the rate of the following reaction:

\[ \text{Zn(s)} + 2\text{HCl(aq)} \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(aq) \]

- [ ] increase \( \text{[HCl]} \)
- [ ] heat up mixture
- [ ] increase surface area of \( \text{Zn(s)} \)
4. Consider the following uncatalyzed reaction which is a one-step (elementary) process:

\[ 2\text{Ce}^{4+} + \text{Ti}^+ \rightarrow 2\text{Ce}^{3+} + \text{Ti}^{3+} \]

When a catalyst is added to the above reaction, the following three step reaction mechanism takes place:

**Step 1.** \( \text{Ce}^{4+} + \text{Mn}^{2+} \rightarrow \text{Ce}^{3+} + \text{Mn}^{3+} \)

**Step 2.** \( \text{Ce}^{4+} + \text{Mn}^{3+} \rightarrow \text{Ce}^{3+} + \text{Mn}^{4+} \)

**Step 3.** \( \text{Mn}^{4+} + \text{Ti}^+ \rightarrow \text{Ti}^{3+} + \text{Mn}^{2+} \)

With reference to the above equation, use collision theory to explain why the catalyzed reaction mechanism is faster than the uncatalyzed reaction.

The catalyzed mechanism involves a rapid series of 2-particle collisions which require less activation energy than the slow 3-particle mechanism. More particles have enough \( E_k \) to react and the rate increases.

5. Consider the following reaction:

\[ \text{CO}_2(g) + \text{NO}_2(g) \rightarrow \text{CO}_2(g) + \text{NO}_2(g) \]

Using collision theory, explain why the rate of the reaction decreases as the reaction proceeds. As the reaction proceeds, the concentration of \( \text{CO}_2 \) and \( \text{NO}_2 \) decrease, this decreases the collision frequency which results in a lower rate of reaction.

6. Consider the following mechanism for an exothermic reaction:

**Step 1.** \( \text{NO}_2(g) + \text{NO}_2(g) \rightarrow \text{N}_2\text{O}_2(g) \) fast

**Step 2.** \( \text{N}_2\text{O}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g) \) slow

Draw a PE diagram to represent the above two-step reaction mechanism and write the net equation to represent the overall reaction.

\[ 2\text{NO}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g) \] slow

7. The uncatalyzed decomposition of methanoic acid, HCOOH, has a \( \Delta H \) of +13 kJ and an activation energy of 88 kJ.

The reaction mechanism for the catalyzed decomposition of methanoic acid is:

**Step 1:** \( \text{HCOOH} + \text{H}^- \rightarrow \text{HCOOH}_2^- \) (Fast)

**Step 2:** \( \text{HCOOH}_2^- \rightarrow \text{HCO}^- + \text{H}_2\text{O} \) (Slow)

**Step 3:** \( \text{HCO}^- \rightarrow \text{H}^- + \text{CO} \) (Fast)

On a graph draw a potential energy diagram for the catalyzed decomposition of methanoic acid. Label the \( \Delta H \) and the activation energy for this reaction.

8. The following equations represent a proposed mechanism for the decomposition of ozone, \( \text{O}_3 \), in the atmosphere.

**Step 1:** \( \text{Cl} + \text{O}_3 \rightarrow \text{ClO} + \text{O}_2 \)

**Step 2:** \( \text{ClO} + \text{O} \rightarrow \text{Cl} + \text{O}_2 \)

(a) Write the equation for the overall reaction.

(b) Identify the catalyst.

(c) Explain how a catalyst increases the rate of a reaction.