Part 1: Looking at reactions

Locate and open the learning object *Looking at reactions*. Select **Start** to display the screen titled *How does a chemical reaction occur?* Read this screen, then select **Start** and adjust the energy slider.

1. Write a description of what happens to the particles in this animation as energy is increased.

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2. Explain in your own words how a reaction occurs.

Select **Next** to display the screen titled *Orientation*. Observe how the particles react as orientation importance is changed.

3. What is the effect on rate of reaction when particle orientation is not important? Explain why.

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4. How is reaction rate affected when particle orientation is important? Explain any difference.

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5. Use collision theory to explain why an increase in energy, and decrease in importance of orientation, leads to an increase in reaction rate.

Select **Next** to display the screen titles *Burning magnesium*.

6. Explain why oxygen and magnesium do not catch alight before they are heated?

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7. Energy is used in the bond-breaking process. What bonds need to be broken?

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8. Energy is released in the bond-making process. Which bonds are being made?

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9. Where does energy go when magnesium reacts with oxygen?

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10. If there is a greater release than input of energy during this reaction, what does this tell you about energy stored in the bonds of the reactants compared with the products?

Select **Next** to display the screen titled *Changing reaction rates*.

11. Describe what happens to the particles as concentration, particle size and temperature are changed. Explain why some variables increase collision rates.

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12. Why does increasing the temperature increase the rate of successful collisions?

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13. Explain what happens to the collision rate as particle size increases.

Select **Next** to display the screen titled *Kinetic energy distribution*.

14. Explain what happens to particles as temperature is increased.

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15. How is temperature related to the kinetic energy of particles?

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16. Does activation energy for a reaction change when temperature is changed? Explain your answer.

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17. Activation energy (E_A) is often described as the reaction 'energy hurdle'. Explain what this means.

Select **Next** to display the screen titled *Energy profile diagrams*.

18. Select **Exothermic reaction**, then move the **Reaction progress** slider across. Describe what happens to the particles at each stage of the reaction animation.

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19. Select **Endothermic reaction**, then move the **Reaction progress** slider across. Describe what happens to the particles at each stage of the reaction animation.

Select **Next** to display the screen titled *Transition state*.

20. What is an activated complex?

21. Why is it possible for an activated complex to form either the products or the reactants?

Select **Next** to display the screen titled *Catalysts*.

22. Describe the difference in activation energies for a reaction with, and without, a catalyst.

23. What does the shaded portion of the energy distribution graph on this screen represent?

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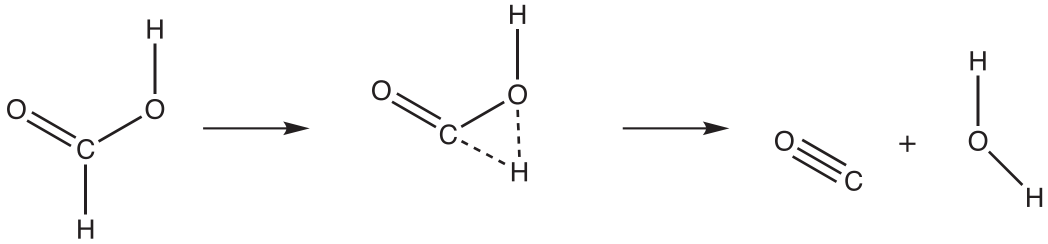
24. Explain how a catalyst increases the rate of reaction?

Figure 1: Energy profile diagram for decomposition of methanoic acid.

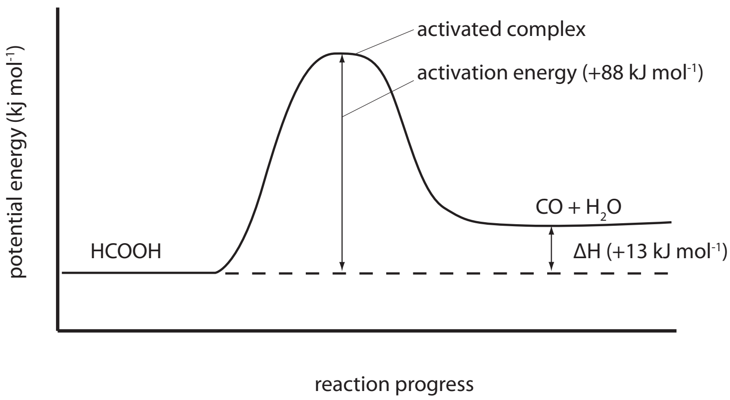
Part 2: How a catalyst can change a reaction pathway

Methanoic acid decomposes to carbon monoxide and water, via an intermediate activated complex.



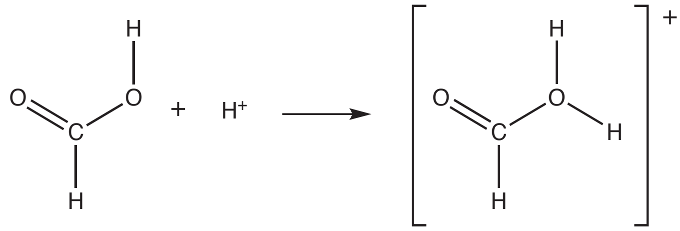
methanoic acid | activated complex | carbon monoxide + water

This reaction proceeds very slowly at room temperature because it has high activation energy. The activation energy is high because a lot of energy is needed to break the bond between carbon and hydrogen before a new bond is formed between oxygen and hydrogen. The activated complex exists only momentarily at the top of the energy profile.

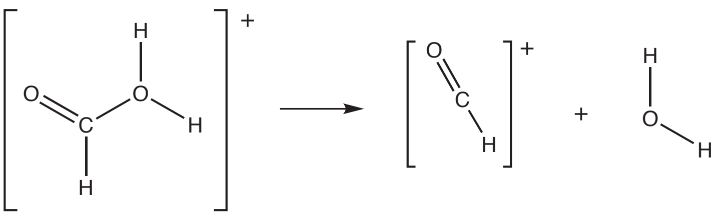


The reaction rate can be increased by adding a catalyst such as hydrogen ions from dilute sulphuric acid. Instead of a single activated state, the reaction now proceeds in three distinct steps that involve two intermediate products and three different activated complexes.

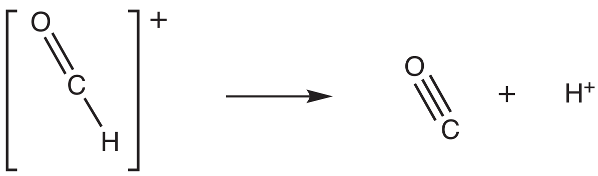
1. A hydrogen ion attaches to a methanoic acid molecule to form the charged molecule, $HCOOH_2^+$.



2. This unstable molecule then splits apart to form a HCO^+ ion and water.



3. In the final step, HCO^+ molecules decompose to form carbon monoxide, which bubbles off, leaving the hydrogen ion catalyst regenerated within the solution.



None of the intermediate steps in the catalysed reaction are as difficult as the single-step decomposition process, so the overall activation energy for the reaction is lowered.

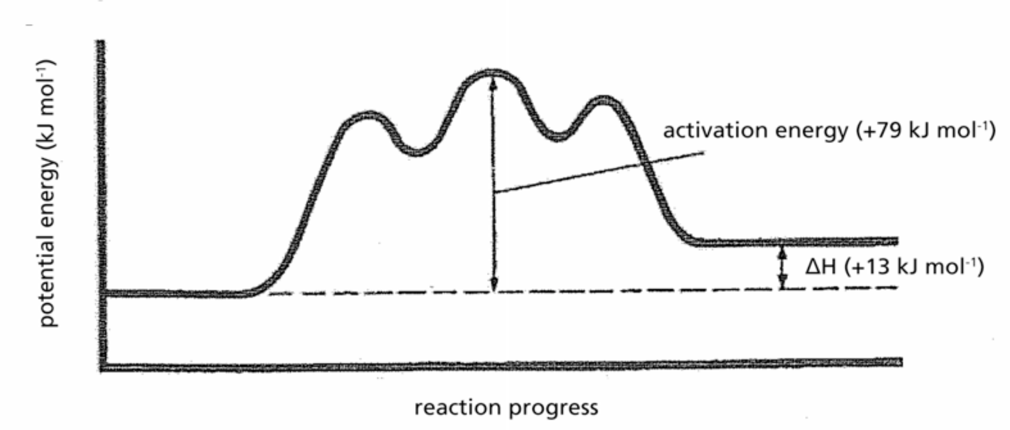


Figure 2: Energy profile diagram for catalysed decomposition of methanoic acid

The energy profile diagram for catalysed decomposition of methanoic acid has three peaks, each corresponding to a step in the reaction described above. The second step has the highest activation energy, but it is still lower than that of the uncatalysed reaction (79 kJ mol^{-1} and 88 kJ mol^{-1} respectively).

The presence of three peaks is not a barrier to reaction, as any molecule with enough kinetic energy to overcome the second peak will have no trouble getting over the first peak on the way up, or the third peak on the way down.

The species $HCOOH_2^+$ and HCO^+ are *not* activated complexes. They are reaction intermediates that exist for a longer time period than the activated complexes. Reaction intermediates are represented by the 'valleys' in the energy profile curve, and activated complexes represented by the peaks.

Questions

1. How does this reaction demonstrate that catalysts take part in a reaction without being used up?

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2. Is this reaction exothermic or endothermic? Explain your answer.

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3. Why is the heat of reaction the same for a catalysed and uncatalysed reaction?

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4. What is the difference between an activated complex and a reaction intermediate?

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5. What might the activated complex, that corresponds to the highest peak of the energy profile curve in Figure 2, look like?