Week 12

# A-Level Chemistry

Homework bundle for this week

本周作业合集

exercises.lol

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A-Level Chemistry Name: _____ Score: _____

1. An effective collision is one that:
   - ☐ has enough energy and the correct orientation, so a reaction happens
   - ☐ has any amount of energy
   - ☐ always leads to a reaction
   - ☐ happens at the wrong angle
2. The activation energy is:
   - ☐ the minimum energy a collision needs to be effective
   - ☐ the energy released by the reaction
   - ☐ the average energy of the molecules
   - ☐ always zero
3. A catalyst speeds up a reaction by:
   - ☐ providing a different route with a lower activation energy
   - ☐ raising the activation energy
   - ☐ being used up in the reaction
   - ☐ increasing the temperature
4. The rate of reaction depends on the:
   - ☐ frequency of effective collisions
   - ☐ colour of the reactants
   - ☐ mass of the container
   - ☐ number of non-effective collisions only
5. On the Boltzmann distribution, the molecules that can react are those:
   - ☐ to the right of the activation energy
   - ☐ to the left of the activation energy
   - ☐ at the peak
   - ☐ at zero energy
6. A small rise in temperature can give a large rise in reaction rate.

   ☐ True ☐ False

7. The rate of reaction can be found from the _____ of a concentration-time graph.

_____

8. Increasing temperature increases the proportion of molecules with energy above the activation energy.

   ☐ True ☐ False

9. A catalyst speeds up a reaction by providing a route with a lower activation energy.

   ☐ True ☐ False

10. A catalyst does NOT change the:

    ☐ enthalpy change (ΔH) of the reaction

    ☐ activation energy

    ☐ rate of reaction

    ☐ reaction mechanism



## 8. Reaction kinetics

**Answers · 答案**

A-Level Chemistry

1. **has enough energy and the correct orientation, so a reaction happens**

   A collision only reacts with sufficient energy and the right orientation.

2. **the minimum energy a collision needs to be effective**

   Only collisions with at least the activation energy can react.

3. **providing a different route with a lower activation energy**

   A catalyst offers a new mechanism with a lower $E_a$, so more molecules can react; it is not used up.

4. **frequency of effective collisions**

   More effective collisions per second means a faster rate.

5. **to the right of the activation energy**

   Molecules with energy greater than $E_a$ (right of the line) have enough energy to react.

6. **True**

   Because the reacting fraction grows sharply with temperature, even a small increase greatly speeds the reaction.

7. **gradient**

   A steeper gradient means a faster rate.

8. **True**

   So more collisions are successful.

9. **True**

   It is not used up.

10. **enthalpy change (ΔH) of the reaction**

    A catalyst lowers $E_a$ and speeds the rate, but the overall ΔH (reactants to products) is unchanged.

A-LEVEL CHEMISTRY • EXERCISE SHEET

# 8.1 Rate of reaction

Name: _____ Class: _____ Date: _____ **Total: 10 marks**

**KEY WORDS** rate of reaction 反应速率 • collision 碰撞 • effective collision 有效碰撞 • collision frequency 碰撞频率 • non-effective collision 无效碰撞

## Objective

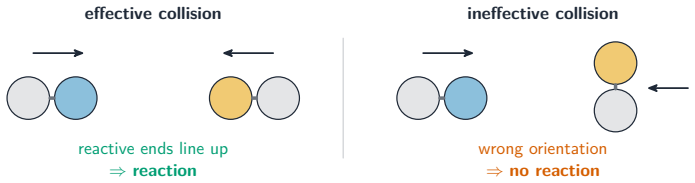
Build the skills to answer exam questions on **rate of reaction** — **collision** 碰撞 theory, **effective collisions** 有效碰撞, and the effect of concentration and pressure.

**You must be able to:**

- explain rate, collision frequency, and effective vs non-effective collisions
- explain the effect of concentration and pressure on rate
- calculate a rate from experimental data

## 1 Worked examples

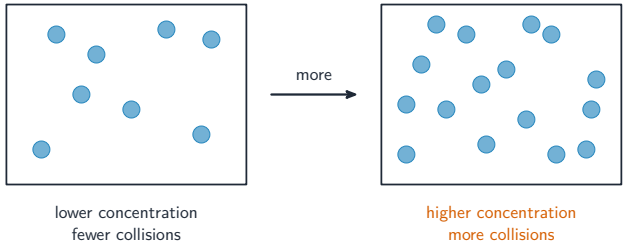
### ■ Collisions and rate




*A collision only reacts with the right orientation and enough energy — an effective collision*

The **rate** is the change in concentration (or amount) per unit time. It depends on the **frequency of effective collisions**.

### ■ Concentration and pressure




*More particles in the same volume collide more often, so the rate rises*

Increasing concentration (or gas pressure) packs particles closer → more effective collisions per second → faster rate.

## 2 Practice

**2.1** State the **two** things a collision needs to be effective. [2]

**2.2** Explain why increasing the concentration of a reactant speeds up a reaction. [2]

## 3 Exam-style questions

**3.1** Increasing the pressure of a gas reaction increases the rate because: [1]

- **A** the particles gain activation energy
- **B** the particles are closer, so collide more often
- **C** the activation energy falls
- **D** the catalyst works better

**3.2** A reaction produces 48 $cm^3$ of gas in 20 s. The mean rate of reaction is: [1]

- **A** 0.42 $cm^3\,s^{-1}$
- **B** 2.4 $cm^3\,s^{-1}$
- **C** 24 $cm^3\,s^{-1}$
- **D** 960 $cm^3\,s^{-1}$

**3.3** A reaction makes carbon dioxide. The volume collected is 30 $cm^3$ at 15 s and 54 $cm^3$ at 45 s.

(a) Calculate the mean rate of reaction between 15 s and 45 s. [2]

(b) Explain why the rate is faster at the start of the reaction than near the end. [2]



## Solutions

**2.1** enough energy (at least the activation energy); the correct orientation/direction.

**2.2** there are more particles in the same volume, so they collide more often — more effective collisions per second.

**3.1 B**. Higher pressure brings particles closer, so they collide more often.

**3.2 B**. $48/20 = 2.4$ $cm^3\,s^{-1}$.

**3.3** (a) $\dfrac{54-30}{45-15} = \dfrac{24}{30} = 0.80$ $cm^3\,s^{-1}$.
(b) at the start the reactant concentration is highest, so collisions are more frequent and the rate is faster; as reactant is used up the rate falls.

**14** 20.0 $cm^3$ of hydrogen peroxide decomposes to water and oxygen in the presence of a suitable catalyst.

160 $cm^3$ of oxygen, measured at room conditions, is produced in 5.00 minutes.

What is the average rate of decomposition of hydrogen peroxide during this reaction period?

**A** $2.22 \times 10^{-5}\,mol\,s^{-1}$

**B** $4.44 \times 10^{-5}\,mol\,s^{-1}$

**C** $1.76 \times 10^{-4}\,mol\,s^{-1}$

**D** $2.67 \times 10^{-3}\,mol\,s^{-1}$





**1** The thiosulfate ion, $S_2O_3^{2-}$, decomposes when an acid is added.

$$S_2O_3^{2-}(aq) + 2H^+(aq) \rightarrow S(s) + SO_2(aq) + H_2O(l)$$

The rate of this reaction can be investigated by measuring how long it takes for the solid sulfur forming to obscure the print on the insert.
You will investigate how the concentration of the thiosulfate ion affects the rate of the reaction.

**Note: A small amount of sulfur dioxide gas may be formed in the experiment. It is very important that you avoid inhaling any fumes. As soon as each experiment is complete, add the reaction mixture to the quenching bath and rinse the beaker thoroughly.**

**FA 1** is 0.100 mol $dm^{-3}$ sodium thiosulfate, $Na_2S_2O_3$.
**FA 2** is 2.00 mol $dm^{-3}$ hydrochloric acid, HC*l*.

**(a) Method**

**Experiment 1**

- Fill a burette with **FA 1**.
- Run 40.00 $cm^3$ of **FA 1** into the 100 $cm^3$ beaker.
- Use the 25 $cm^3$ measuring cylinder to measure 10.0 $cm^3$ of **FA 2**.
- Add the **FA 2** to the **FA 1** in the beaker and start timing immediately.
- Stir the mixture once and place the beaker on the printed insert.
- View the printing on the insert from above through the solution.
- Stop timing when the print on the insert becomes obscured.
- Record this reaction time to the nearest second in the space for results.
- Empty the contents of the beaker into the quenching bath.
- Rinse and dry the beaker so it is ready to use in **Experiment 2**.

**Experiment 2**

- Refill the burette with **FA 1**.
- Fill the second burette with distilled water.
- Run 20.00 $cm^3$ of **FA 1** into the 100 $cm^3$ beaker.
- Run 20.00 $cm^3$ of distilled water into the same beaker.
- Use the 25 $cm^3$ measuring cylinder to measure 10.0 $cm^3$ of **FA 2**.
- Add the **FA 2** to the **FA 1** in the beaker and start timing immediately.
- Stir the mixture once and place the beaker on the printed insert.
- View the printing on the insert from above through the solution.
- Stop timing when the print on the insert becomes obscured.
- Record this reaction time to the nearest second in the space for results.
- Empty the contents of the beaker into the quenching bath.
- Rinse and dry the beaker so it is ready to use in the next experiment.

**Experiments 3–5**

- Carry out three further experiments to investigate how using different volumes of **FA 1** affects the reaction time.

**Note**: the combined volumes of **FA 1** and distilled water must always be 40.00 $cm^3$.
Do not use a volume of **FA 1** that is less than 15.00 $cm^3$.

Record all your results in a table. You should include the volume of **FA 1**, the volume of distilled water, the reaction time and the reaction rate for each of your five experiments.

The rate of reaction can be calculated using the following formula.

$$\text{rate} = \frac{1000}{\text{reaction time}}$$

**Results**

| | |
|---|---|
| I | |
| II | |
| III | |
| IV | |
| V | |
| VI | |
| VII | |
| VIII | |

[8]

**(b) (i)** On the grid in Fig. 1.1, plot the rate (*y*-axis) against the volume of **FA 1** (*x*-axis).
Start each axis at the origin (0, 0).

Ring any anomalous points. Draw a line of best fit.





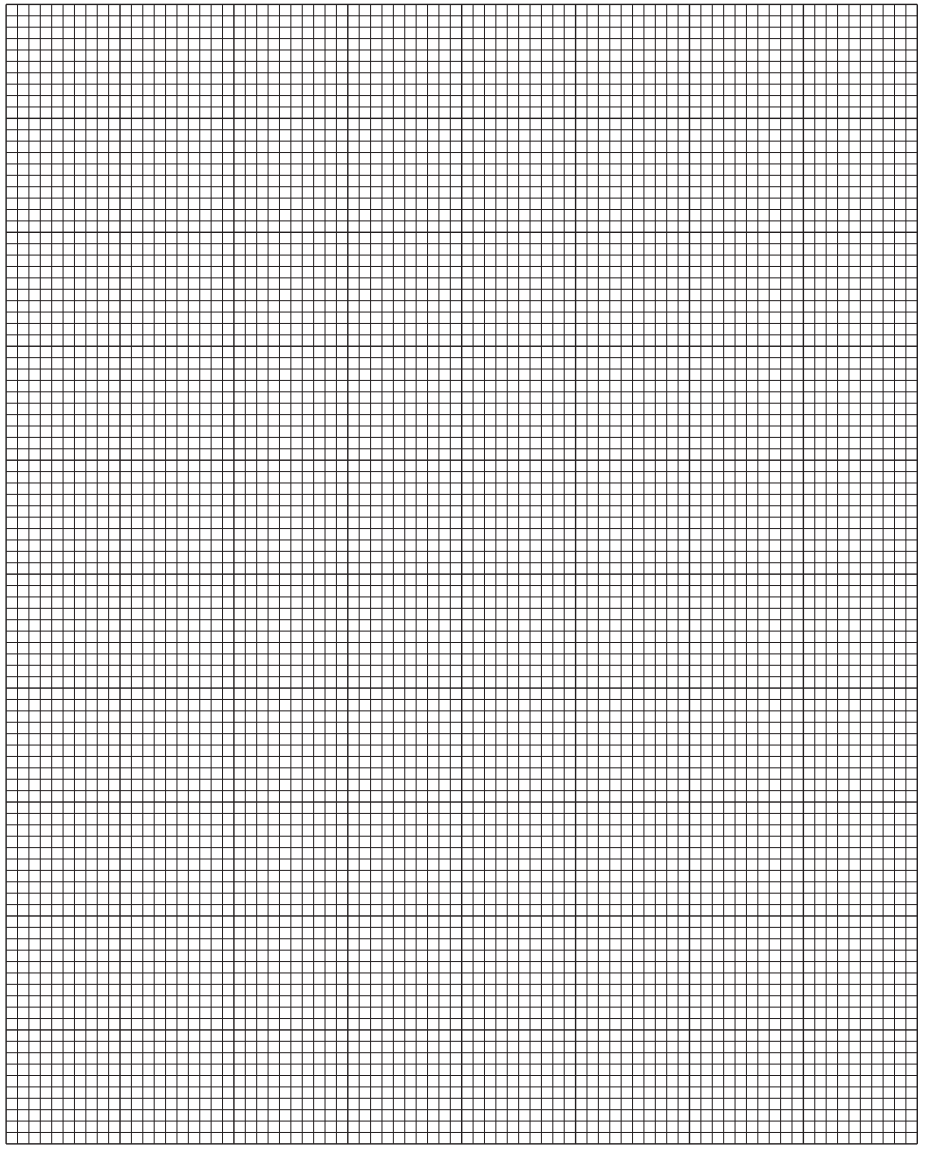


**Fig. 1.1**

| | |
|---|---|
| I | |
| II | |
| III | |

[3]

**(ii)** Use your graph in Fig. 1.1 to determine the time it would take for the print to be obscured if 7.50 $cm^3$ of **FA 1**, 32.50 $cm^3$ of water and 10.0 $cm^3$ of **FA 2** had been used.
Show clearly on the graph how you worked out your answer.

time for printing to be obscured = ..... s [2]

**(c)** A student carries out the same method as given in **(a)** but using a different concentration of acid. The student's calculated values for the rate are plotted on the grid in Fig. 1.2.

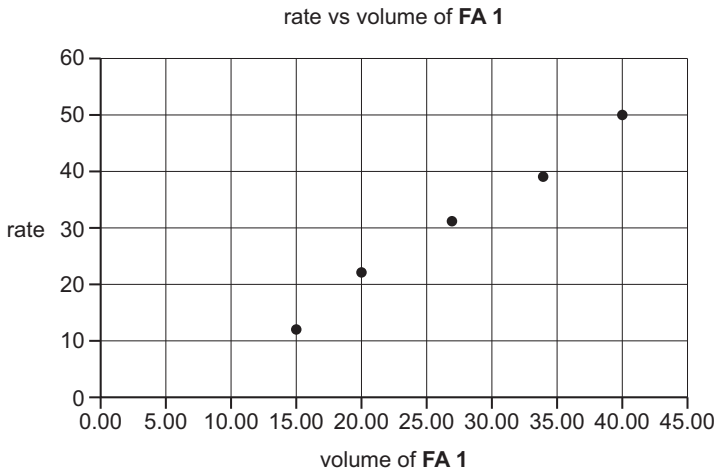



**Fig. 1.2**

State whether the student's results show that the rate is directly proportional to the volume of **FA 1** used.
Explain your answer.

.....

..... [2]

[Total: 15]



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# 8.2 Effect of temperature on reaction rates and activation energy

Name: _____ Class: _____ Date: _____ **Total: 10 marks**

**KEY WORDS** activation energy 活化能 • boltzmann distribution 玻尔兹曼分布 • rate of reaction 反应速率 • effective collision 有效碰撞 • collide 碰撞

## Objective

Build the skills to answer exam questions on **temperature and activation energy** — **activation energy** 活化能, the **Boltzmann distribution** 玻尔兹曼分布, and why temperature has a large effect on rate.

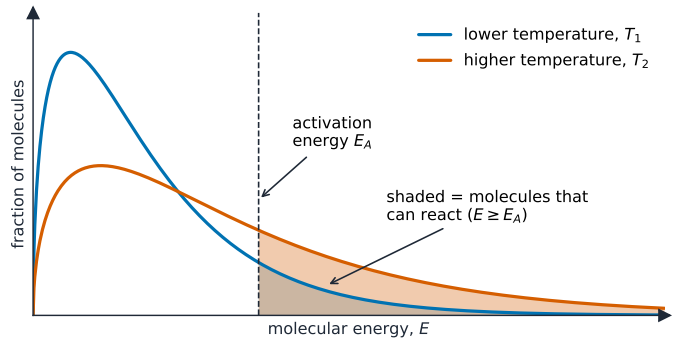
**You must be able to:**

- define activation energy and sketch the Boltzmann distribution
- use the Boltzmann distribution to explain the significance of $E_A$
- explain the effect of temperature in terms of the distribution and effective collisions

## 1 Worked examples

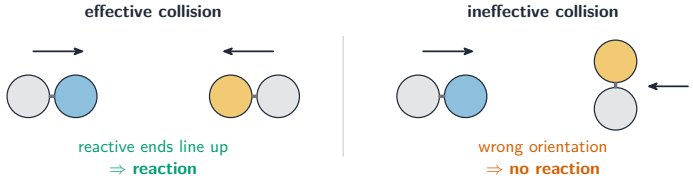
### ■ Activation energy and the Boltzmann distribution

The **activation energy** $E_A$ is the minimum energy a collision needs to be effective.




higher $T$ ⇒ more molecules with $E \geq E_A$

*The curve starts at the origin, peaks, then tails off; only molecules to the right of $E_A$ can react. At higher $T$ the curve spreads right, so a larger fraction can react*




*Higher temperature gives more effective collisions*

### ■ Why temperature has a big effect

Raising the temperature:

- spreads the curve to the right, so a **much larger fraction** of molecules exceed $E_A$ (the bigger effect);
- makes molecules move faster, so they collide more often.

So a small temperature rise gives a large rate rise.

## 2 Practice

**2.1** Define **activation energy**. [1]



**2.2** On a Boltzmann distribution, state which molecules are able to react. [1]

## 3 Exam-style questions

**3.1** When the temperature is increased, the main reason the rate rises is: [1]

- **A** the activation energy falls
- **B** a much larger fraction of molecules have energy $\geq E_A$
- **C** the molecules get bigger
- **D** the enthalpy change increases

**3.2** The total area under a Boltzmann distribution curve represents: [1]

- **A** the activation energy
- **B** the total number of molecules
- **C** the rate of reaction
- **D** the temperature

**3.3** Sketch a Boltzmann distribution curve. Mark the activation energy $E_A$ and shade the molecules that can react. [3]

**3.4** Explain, using the Boltzmann distribution, why a 10 °C rise in temperature can roughly double the rate of a reaction. [3]

## Solutions

**2.1** the minimum energy a collision needs to be effective (to react).

**2.2** those with energy greater than (to the right of) the activation energy.

**3.1 B**. A higher temperature gives a larger fraction of molecules with energy $\geq E_A$.

**3.2 B**. The area under the curve is the total number of molecules.

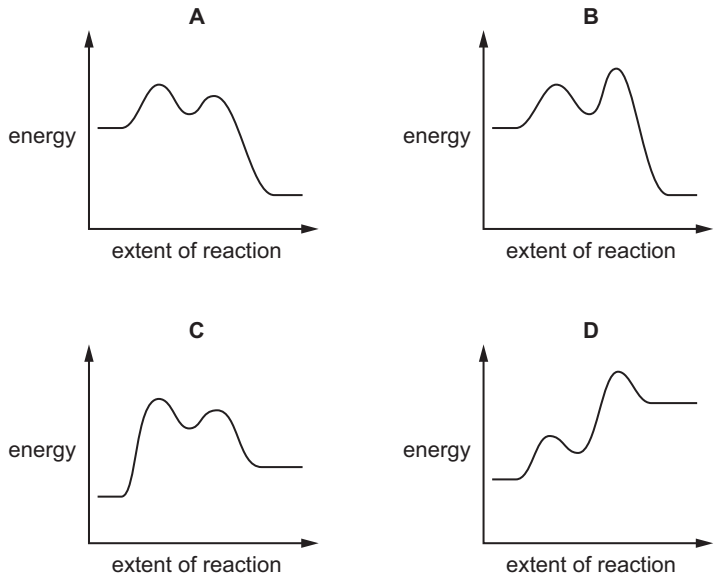
**3.3** curve from the origin rising to a peak then tailing off, not touching the x-axis; $E_A$ marked on the energy axis to the right of the peak; area to the right of $E_A$ shaded.

**3.4** at higher temperature the curve spreads to the right; the fraction of molecules with energy $\geq E_A$ increases greatly; so many more collisions are effective each second, roughly doubling the rate.





**15** Which reaction pathway diagram shows an endothermic reaction that occurs in two steps and in which the second step of the reaction is likely to be faster than the first?






**15** A mixture of gases reacts faster as its temperature increases.

Which row explains this?

| | the activation energy remains unchanged | more particles have energy equal to or above the activation energy | there is an increase in the rate of successful collisions |
|---|---|---|---|
| **A** | false | false | false |
| **B** | true | true | true |
| **C** | false | true | true |
| **D** | true | false | false |

A-LEVEL CHEMISTRY • EXERCISE SHEET

# 8.3 Homogeneous and heterogeneous catalysts

Name: _____ Class: _____ Date: _____ **Total: 13 marks**

**KEY WORDS** catalyst 催化剂 • heterogeneous catalyst 多相催化剂 • mechanism 反应机理 • homogeneous 均相 • reaction pathway diagram 反应路径图

## Objective

Build the skills to answer exam questions on **catalysts** 催化剂 — how a catalyst works, and the difference between **homogeneous** 均相 and **heterogeneous** 多相 catalysts.

**You must be able to:**

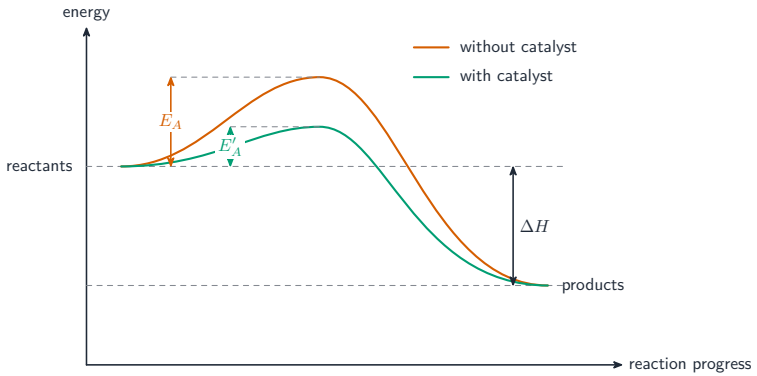
- explain that a catalyst gives a different mechanism of lower activation energy
- explain the catalytic effect using the Boltzmann distribution
- construct and interpret reaction pathway diagrams with and without a catalyst

## 1 Worked examples

### ■ How a catalyst works

A **catalyst** speeds up a reaction without being used up. It provides a different **mechanism** 反应机理 with a **lower activation energy**.

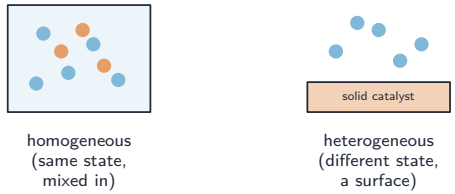





*The catalysed route has a lower $E_A$; $\Delta H$ is unchanged*

On the Boltzmann distribution, lowering $E_A$ moves the line left, so more molecules can react → faster rate.

### ■ Homogeneous vs heterogeneous




*Homogeneous: same state as the reactants (e.g. acid in solution). Heterogeneous: a different state (e.g. solid iron with gases)*

## 2 Practice

**2.1** State what a catalyst does and what happens to it during the reaction. [2]

**2.2** State the difference between a homogeneous and a heterogeneous catalyst. [2]

## 3 Exam-style questions

**3.1** A catalyst increases the rate of a reaction by: [1]

- **A** increasing the activation energy
- **B** providing a route with a lower activation energy
- **C** making the reaction more exothermic
- **D** raising the temperature

**3.2** Which property of a reaction is **not** changed by adding a catalyst? [1]

- **A** the rate
- **B** the activation energy
- **C** the enthalpy change, $\Delta H$
- **D** the mechanism

**3.3** (a) Explain, using the Boltzmann distribution, why a catalyst speeds up a reaction. [2]

(b) Sketch a reaction pathway diagram showing the catalysed and uncatalysed routes for an exothermic reaction. [3]

**3.4** Iron is used in the Haber process. State whether it is a homogeneous or heterogeneous catalyst, and explain your answer. [2]



## Solutions

**2.1** it speeds up the reaction; it is not used up / is regenerated.

**2.2** homogeneous — in the same physical state as the reactants; heterogeneous — in a different state.

**3.1 B**. A catalyst provides a route with lower activation energy.

**3.2 C**. The enthalpy change is unchanged by a catalyst.

**3.3** (a) the catalyst lowers the activation energy; so a larger fraction of molecules now exceed $E_A$ and can react — more effective collisions per second.
(b) reactants higher than products (exothermic); two activation-energy humps, the catalysed one lower; same $\Delta H$ marked between reactants and products.

**3.4** heterogeneous; the iron is a solid while the reactants ($N_2$ and $H_2$) are gases — a different physical state.

**15** The reversible reaction between methanol and ethanoic acid liquids is catalysed by adding a small volume of concentrated sulfuric acid.

Two statements about this reaction are listed.

1 The sulfuric acid is a homogeneous catalyst.

2 The sulfuric acid lowers the activation energy of the reverse reaction.

Which statements are correct?

**A** both 1 and 2

**B** 1 only

**C** 2 only

**D** neither 1 nor 2



**14** When two aqueous solutions are mixed, the reaction between them is very slow.

An effective catalyst is added to the mixture without any change in temperature.

Which statement about this catalyst is correct?

**A** It increases the average energy of the reactant particles.

**B** It changes the distribution of energies of the reactant particles.

**C** It allows a greater number of reactant particles to react per unit time.

**D** It increases the number of reactant particles with the most probable energy.





**2** Aluminium oxide, $Al_2O_3$, and phosphorus(V) oxide, $P_4O_{10}$, are both used as reagents and catalysts.

**(a)** The melting point of $Al_2O_3$ is 2072 °C. The melting point of $P_4O_{10}$ is 340 °C.

Explain the difference in the melting points of these two compounds.

.....

.....

.....

..... [3]

**(b)** A 5.00 $dm^3$ sealed flask contains 0.400 mol of CO(g) and 0.800 mol of $H_2$(g) and an $Al_2O_3$ catalyst. The flask is heated to a temperature of 290 °C and allowed to reach equilibrium. Equation 1 shows the reaction.

equation 1 $\quad CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$

The equilibrium constant, $K_c$, of equation 1 is given.

$$K_c = \frac{[CH_3OH]}{[CO][H_2]^2}$$

**(i)** State the units of $K_c$.

..... [1]

**(ii)** The equilibrium mixture contains 0.280 mol of $CH_3OH$(g).

Calculate the value of $K_c$.

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

value of $K_c$ = ..... [3]

**(iii)** State and explain the effect, if any, on the value of $K_c$ when the overall pressure in the sealed flask is increased.

..... .....

..... [1]

**(c)** $P_4O_{10}$ catalyses the reversible reaction of CO with $H_2$ to form $CH_3OH$.

equation 1 $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$

$P_4O_{10}$ then acts as a dehydrating agent, causing $CH_3OH$ to form $CH_3OCH_3$.

**(i)** Explain how the presence of a catalyst affects a chemical reaction.

.....

.....

..... [1]

**(ii)** Construct an equation for the dehydration reaction of $CH_3OH$ to form $CH_3OCH_3$.

..... [1]

**(iii)** Write an equation to show the reaction of $P_4O_{10}$ with an excess of water.

..... [1]

[Total: 11]