

# Mastering Chemistry

- Book 6
- Topic 13 Industrial Chemistry



## Content

- ➔ 44.1 The activation energy barrier
- ➔ 44.2 Potential energy profiles
- ➔ 44.3 The kinetic energies of moving molecules in a gas
- ➔ 44.4 Effect of temperature on the rate of a gaseous reaction
- ➔ 44.5 Effect of temperature on rate constants

Continued on next page ➔



# Content

- ➔ 44.6 What does a catalyst do?
- ➔ 44.7 The Maxwell-Boltzmann distribution curve and catalyst
- ➔ 44.8 Catalysts in industry
- ➔ 44.9 Enzymes
- ➔ Key terms
- ➔ Summary
- ➔ Unit Exercise



## 44.1 The activation energy barrier (p.45)

- ◆ Most reactions are exothermic, and very few are endothermic. If a reaction is highly exothermic, it means the chemical energy of the products is much less than the chemical energy of the reactants.
- ◆ Then you may think the reaction will take place very rapidly. This, however, is not necessarily the case.
- ◆ For example, a mixture of methane and oxygen shows no sign of reaction at room temperature and pressure even though the standard enthalpy change of complete combustion of methane is  $-890 \text{ kJ mol}^{-1}$ .



## 44.1 The activation energy barrier (p.45)

- ◆ The enthalpy change of reaction only indicates the heat change in a process. It does not tell the fact that in order to start off the process going from reactants to products, energy barrier must be overcome.
- ◆ The enthalpy change of a reaction does not indicate the rate of the reaction.



## 44.1 The activation energy barrier (p.45)

### Collision theory

- ◆ Consider this reaction:  $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightarrow 2\text{HI}(\text{g}) \quad \Delta H < 0$
- ◆ In order for one  $\text{H}_2$  molecule to react with one  $\text{I}_2$  molecule, the two molecules must first collide with each other. However, not all collisions between reactant molecules will result in a reaction. There are two requirements for a reaction to occur:
  - For a reaction to take place, the two colliding molecules must have a certain minimum kinetic energy, the **activation energy** (活化能).
  - The two molecules must collide in a correct orientation.



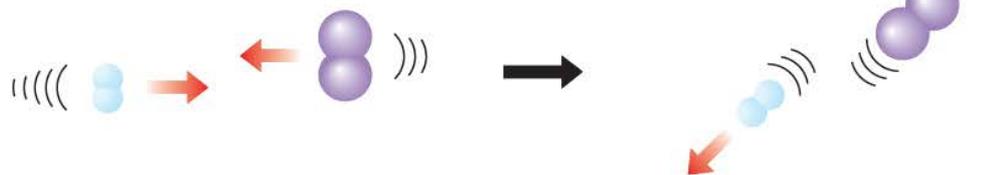
# 44.1 The activation energy barrier (p.45)

two molecules collide with a total kinetic energy less than the activation energy



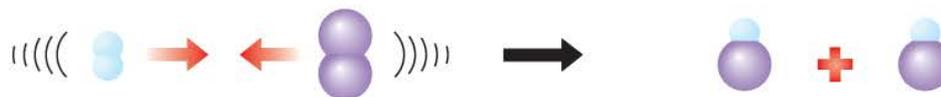
molecules bounce apart — no reaction

two molecules collide in a wrong orientation



molecules bounce apart — no reaction

two molecules collide with a total kinetic energy equal to or greater than the activation energy and in the right orientation



reaction takes place

key:

 hydrogen molecule

 iodine molecule



## 44.1 The activation energy barrier (p.45)

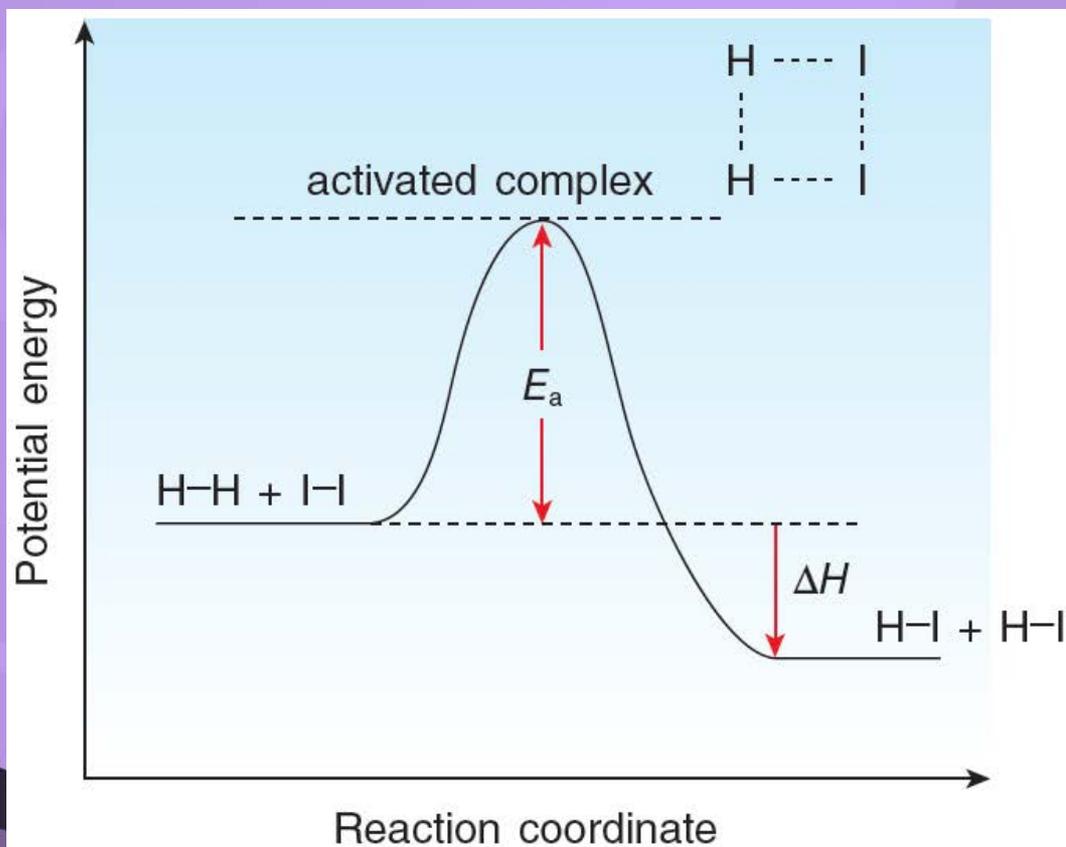
- ◆ All chemical reactions result in the breaking of some bonds (which absorbs energy) and the forming of new ones (which releases energy).
- ◆ Energy is used to overcome the repulsion between electron clouds of the colliding particles and to start rearranging the bonds to form the product(s).

**Activation energy is the minimum kinetic energy that colliding particles must have for a collision to result in a chemical reaction.**



## 44.2 Potential energy profiles (p.47)

- ◆ A **potential energy profile** (勢能能線圖) can be used to show the potential energy changes occurring during a chemical reaction.



key:

$E_a$  = activation energy for the reaction

$\Delta H$  = enthalpy change of reaction



## 44.2 Potential energy profiles (p.47)

- ◆ As two reactant molecules with enough kinetic energies collide with a correct orientation, their kinetic energy is converted to potential energy. The H–H bond and I–I bond begin to break and the H–I bonds begin to form.
- ◆ As bond breaking and bond formation take place, the interacting molecules are temporarily in a highly energetic and unstable state. This state which occurs at the point of maximum potential energy of the profile is known as the **transition state (過渡態)** or **activated complex (活化複體)**.



**Energy profile for a reaction**  
[Ref.](#)



## 44.2 Potential energy profiles (p.47)

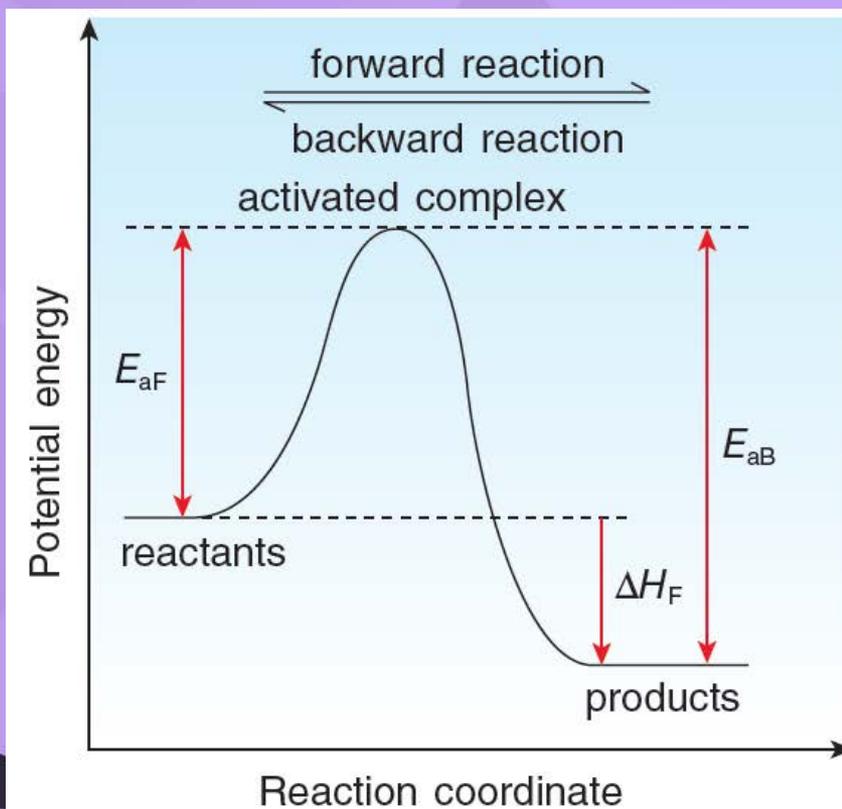
- ◆ As the H–I bonds continue to form, the potential energy of the molecules falls.
- ◆ In general, reactions with low activation energies take place very rapidly, because the energy needed to break the bonds is readily available from the surroundings. Reaction with very high activation energies take place extremely slowly or do not take place at all.



## 44.2 Potential energy profiles (p.47)

Potential energy profile for a reversible reaction

- ◆ The forward reaction and the backward reaction have the same activated complex.



key:

$E_{aF}$  = activation energy for the forward reaction

$E_{aB}$  = activation energy for the backward reaction

$\Delta H_F$  = enthalpy change of the forward reaction



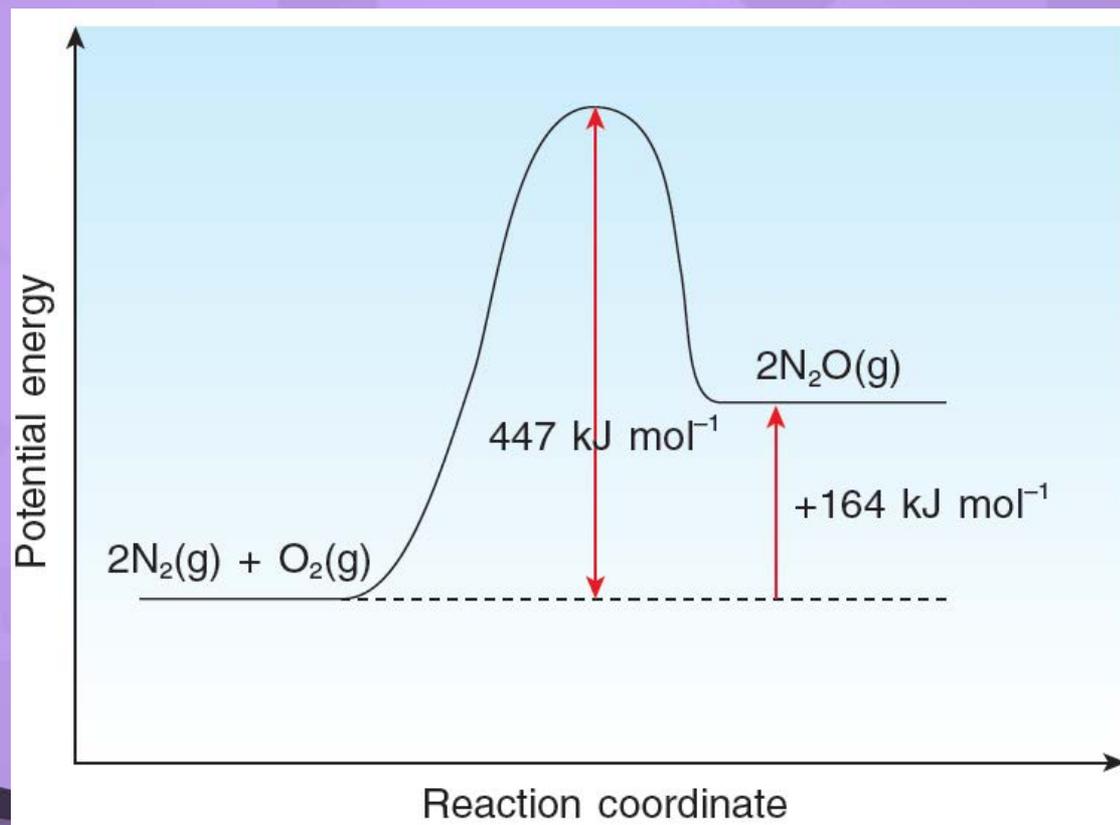
## 44.2 Potential energy profiles (p.47)

### Practice 44.1

Under certain conditions nitrogen reacts with oxygen to make  $\text{N}_2\text{O}$ .



The potential energy profile for this reaction is shown below.





## 44.2 Potential energy profiles (p.47)

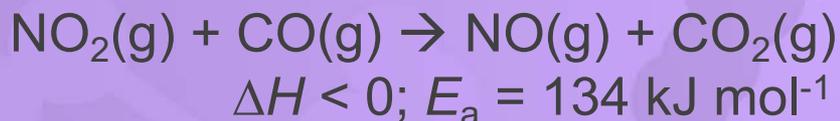
a) What is the enthalpy change of formation of  $\text{N}_2\text{O}(\text{g})$ ?

$$\frac{+164}{2} \text{ kJ mol}^{-1} = +82 \text{ kJ mol}^{-1}$$

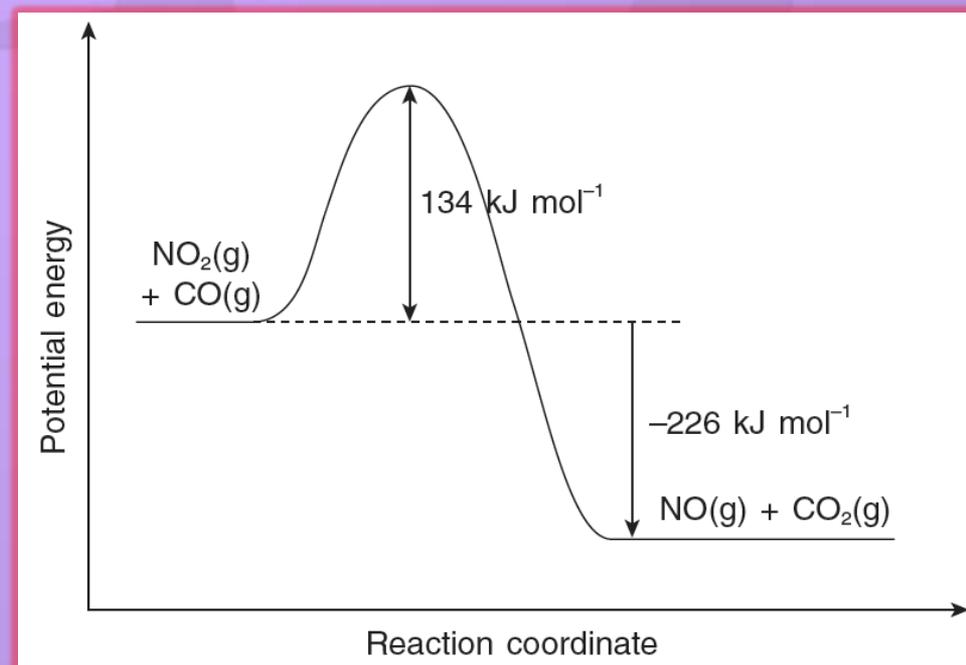
b) What is the activation energy for the backward reaction?

$$(447 - 164) \text{ kJ mol}^{-1} = 283 \text{ kJ mol}^{-1}$$

2 Given the data for the reaction:



Draw a potential energy profile for the reaction.



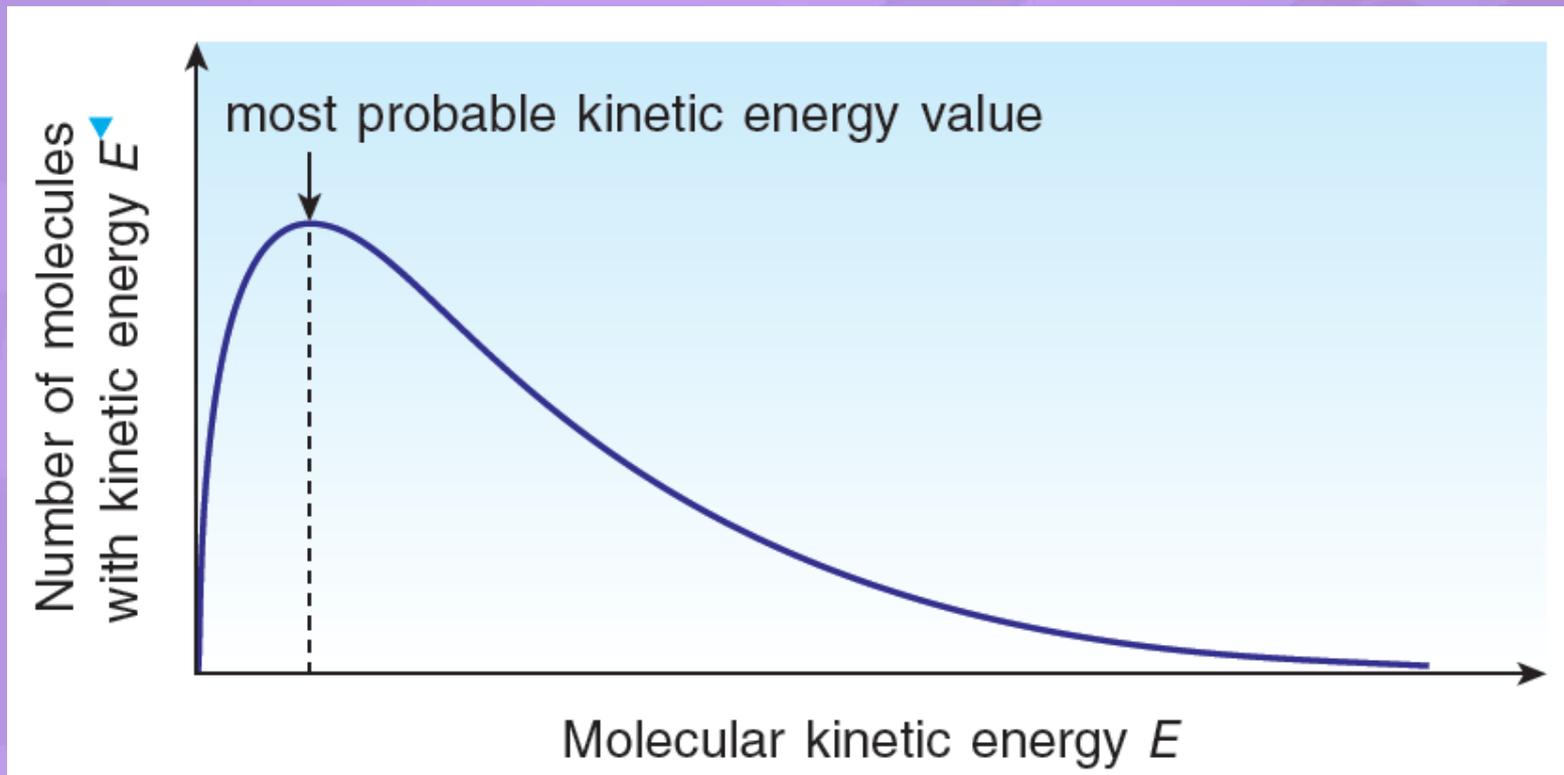


## 44.3 The kinetic energies of moving molecules in a gas (p.49)

- ◆ For a gas, the molecules move around inside a container. They collide with each other and with the walls of the container. Some molecules move slowly with low kinetic energy. Some molecules move fast with high kinetic energy. Most molecules move at an average speed and have an average value of kinetic energy.
- ◆ The spread of molecular kinetic energy in a gas is called the **Maxwell-Boltzmann distribution** (麥克斯韋-波爾茲曼分佈).



## 44.3 The kinetic energies of moving molecules in a gas (p.49)





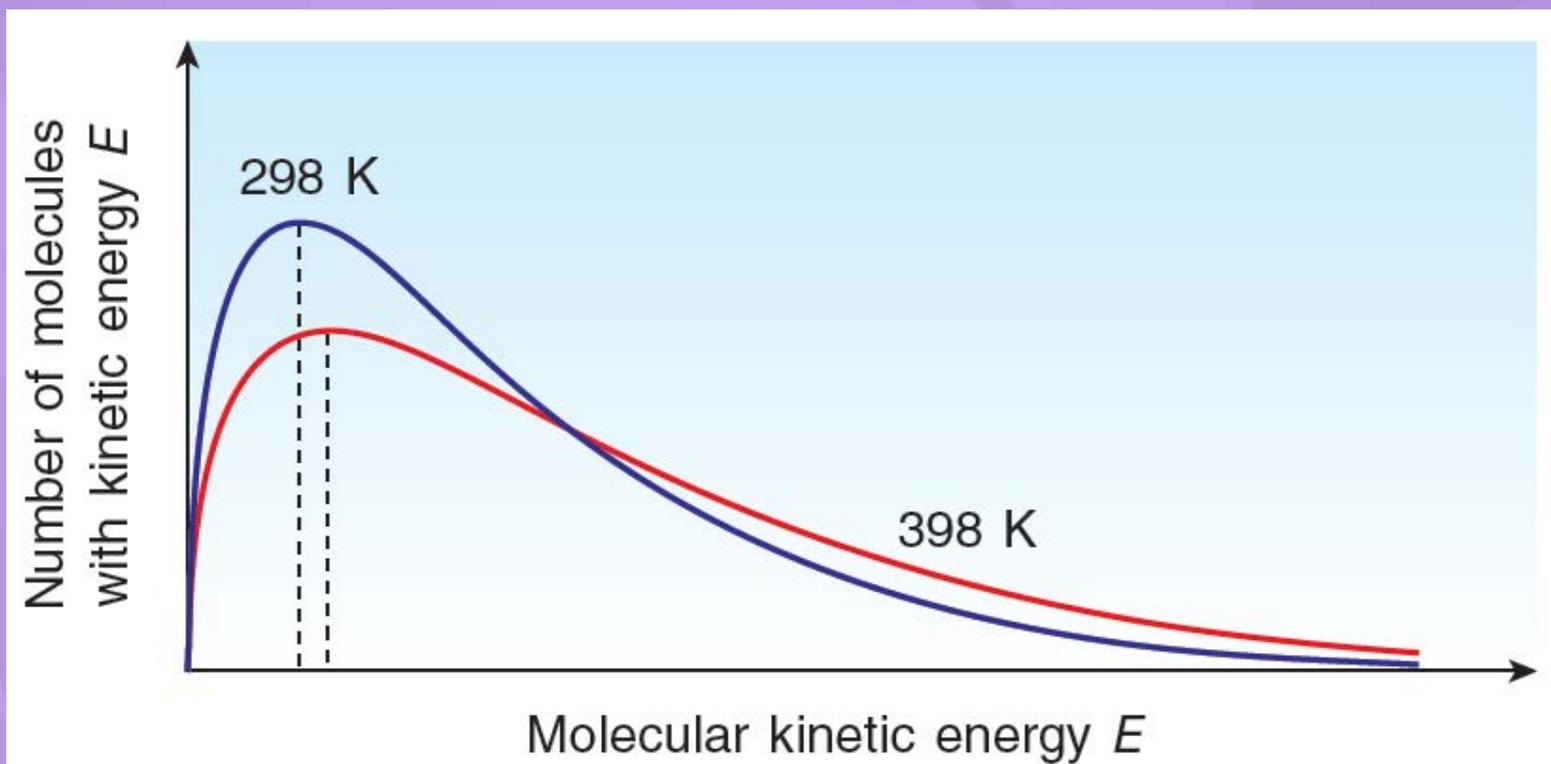
## 44.3 The kinetic energies of moving molecules in a gas (p.49)

- ◆ The area under the curve is equal to the total number of molecules in the gas.
- ◆ Each molecule in the system has kinetic energy greater than zero—the curve starts at the origin.
- ◆ There is no maximum kinetic energy for a molecule — the curve does not meet the kinetic energy axis.
- ◆ The highest point on the curve shows the **most probable kinetic energy** (最大機率動能).



## 44.3 The kinetic energies of moving molecules in a gas (p.49)

Effect of temperature on Maxwell-Boltzmann distribution





## 44.3 The kinetic energies of moving molecules in a gas (p.49)

- ◆ Both curves start at the origin and finish by approaching the x-axis.
- ◆ The area under both curves is the same (the number of molecules is the same).
- ◆ At the higher temperature, the curve is flatter and its maximum has moved to the right — there are more molecules with higher kinetic energies.

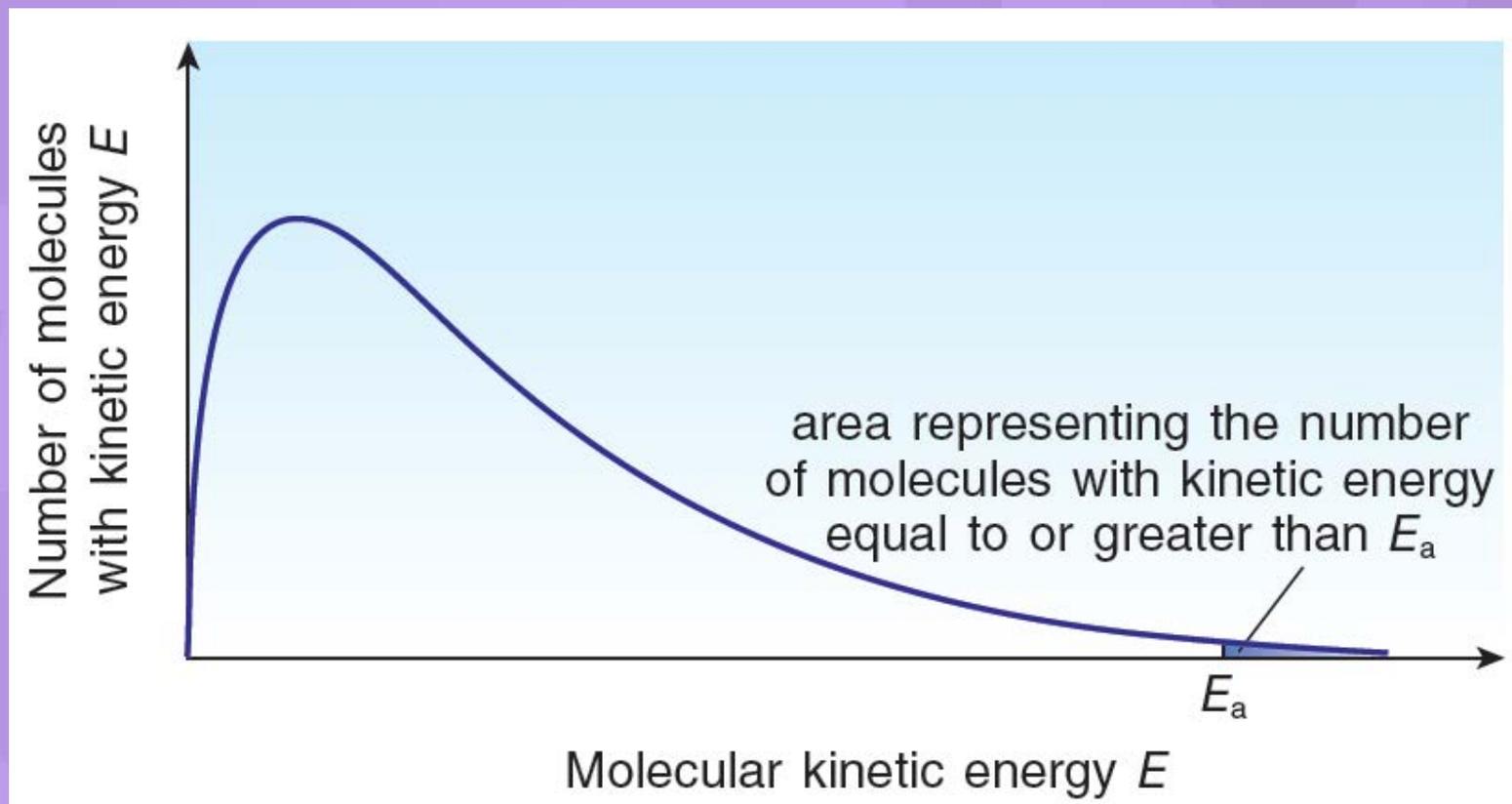


## 44.4 Effect of temperature on the rate of a gaseous reaction (p.51)

- ◆ For a gaseous reaction to take place, only those colliding molecules with kinetic energy equal to or greater than the activation energy can react.
- ◆ If you mark the Maxwell-Boltzmann distribution curve with a line,  $E_a$ , which represents the activation energy, you can see from the shaded area that only a small fraction of the molecules have sufficient energy to react. For a reaction that proceeds steadily at room temperature, only about 1 in  $10^{12}$  molecules possesses this energy.



## 44.4 Effect of temperature on the rate of a gaseous reaction (p.51)





## 44.4 Effect of temperature on the rate of a gaseous reaction (p.51)

- ◆ When you raise the temperature of a gaseous reaction mixture,
  - the average kinetic energy of the molecules increases;
  - molecules move around more quickly;
  - the frequency of collisions increases.

Hence there are more effective collisions in a unit volume per unit time. The rate of reaction increases.

- ◆ However, this has only a very minor effect on the rate of reaction and cannot account for the rapid increase in reaction rate as the temperature increases.



## 44.4 Effect of temperature on the rate of a gaseous reaction (p.51)

The behaviour of gases is explained by the kinetic theory of gases. According to the kinetic theory, collision frequency is proportional to the square root of the absolute temperature.

$$\text{collision frequency} \propto \sqrt{T}$$

Suppose the temperature is increased from 298 K to 308 K, the ratio of the collision frequencies at the two temperatures equals

$$\sqrt{\frac{308}{298}}, \text{ i.e. } 1.017.$$

These calculations show that the collision frequency increases only 1.017 times for a 10-degree rise in temperature.

In fact, a 10-degree rise in temperature nearly doubles the reaction rate. This indicates that the increase in reaction rate with temperature rise is not due to the increase in collision frequency alone.



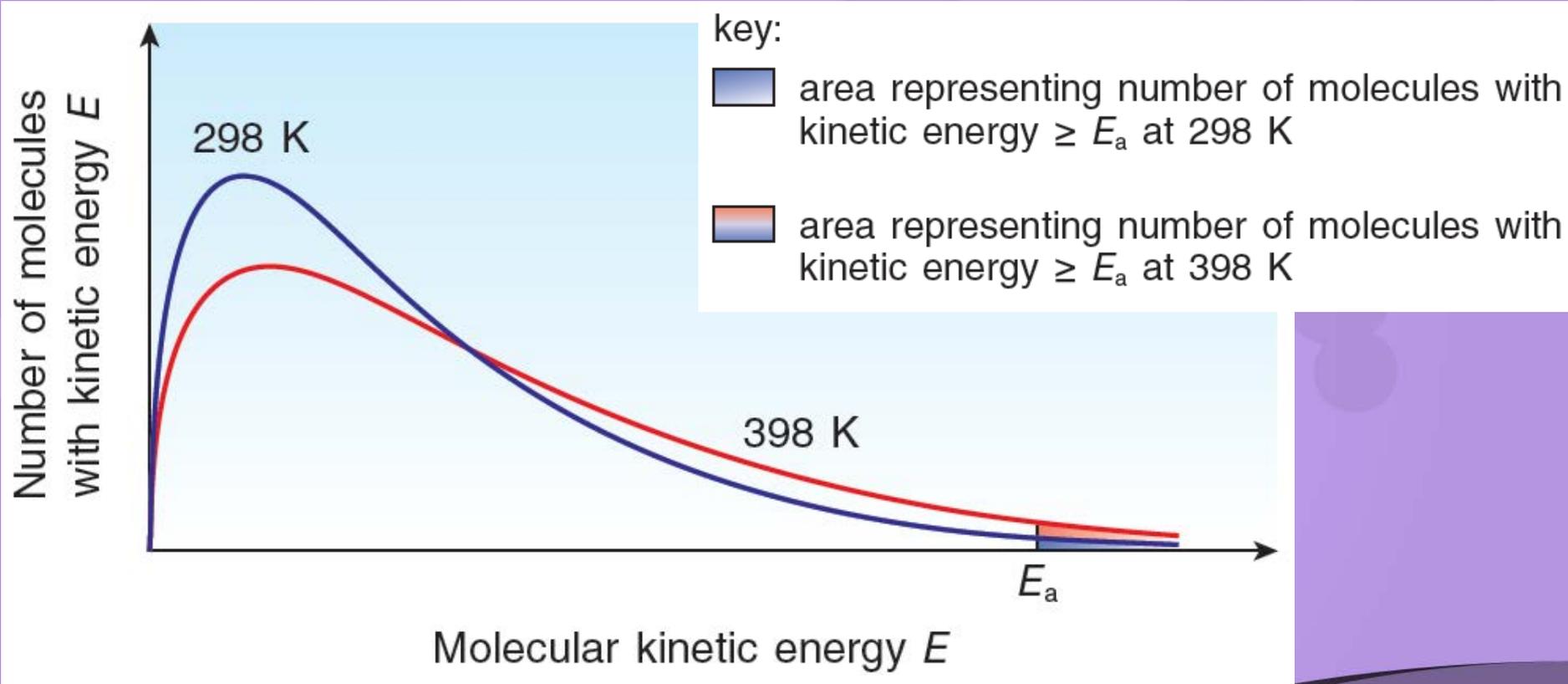
## 44.4 Effect of temperature on the rate of a gaseous reaction (p.51)

- ◆ When the temperature is increased,
  - the area under the curve to the right of the line  $E_a$  increases;
  - this means that at a higher temperature, a greater fraction of molecules has kinetic energy equal to or greater than the activation energy.

Hence a higher percentage of molecules with sufficient energy to collide and react. The rate of reaction increases rapidly.



## 44.4 Effect of temperature on the rate of a gaseous reaction (p.51)





## 44.4 Effect of temperature on the rate of a gaseous reaction (p.51)

In summary, a temperature rise

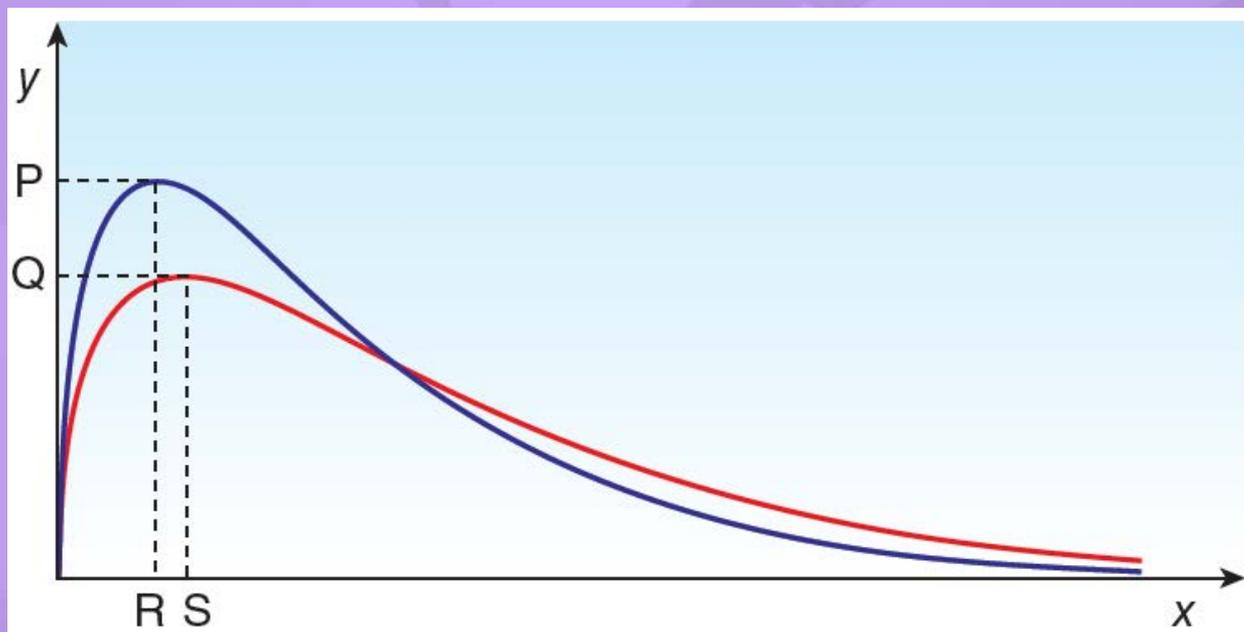
- ◆ increases the average kinetic energy of the molecules, causing the collision frequency to increase, and hence effective collisions among molecules occur more frequently (accounts for a small increase in the rate of reaction).
- ◆ increases the fraction of molecules with kinetic energy equal to or greater than the activation energy, and hence a higher percentage of molecules with sufficient energy to collide and react (accounts for a rapid increase in the rate of reaction).



## 44.4 Effect of temperature on the rate of a gaseous reaction (p.51)

### Practice 44.2

The Maxwell-Boltzmann distribution curves of molecular kinetic energies of a gas at two different temperatures are shown below.





## 44.4 Effect of temperature on the rate of a gaseous reaction (p.51)

- a) What do the axes,  $x$  and  $y$ , represent respectively?  
 **$x$ -axis: molecular kinetic energy  $E$**   
 **$y$ -axis: number of molecules with kinetic energy  $E$**
- b) What does the area under a Maxwell-Boltzmann distribution curve represent? **The total number of molecules in the gas**
- c) Which letter, P, Q, R or S, represents the most probable kinetic energy of the molecules at the lower temperature? **R**
- d) With reference to the curves, explain why a decrease in temperature decreases the reaction rate of a gaseous reaction.  
**Upon a temperature decrease, the reactant particles have less kinetic energy.**  
**The collision frequency among molecules decreases, and hence effective collisions among molecules occur less frequently.**  
**The fraction of molecules with kinetic energy equal to or greater than the activation energy decreases. Thus, fewer molecules with sufficient energy to collide and react.**



## 44.5 Effect of temperature on rate constants (p.53)

- ◆ As you have seen, the rate equation for a reaction between two reactants A and B looks like this:

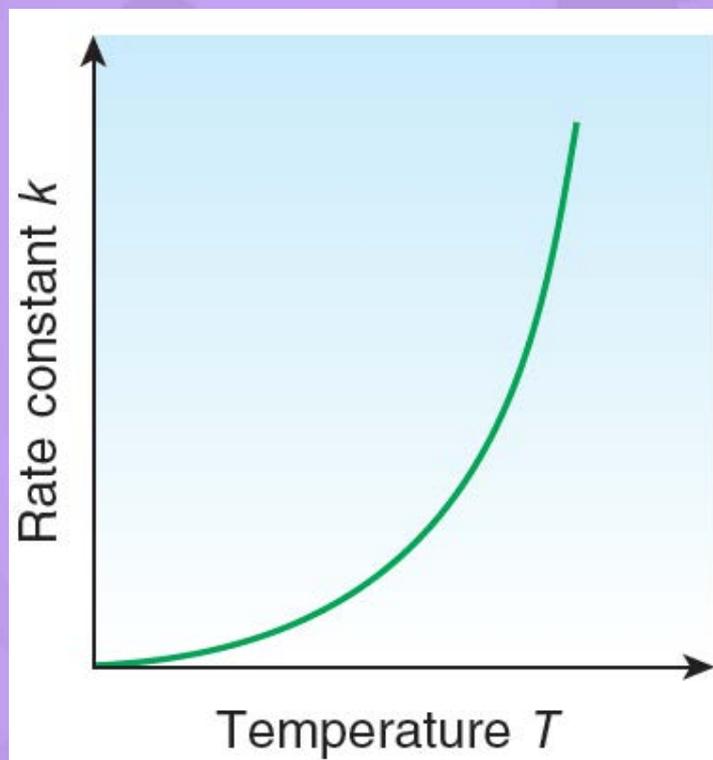
$$\text{rate} = k[A]^p[B]^q$$

- ◆ The equation shows the effect of changing the concentrations of the reactants on the rate of the reaction. However, other variables such as temperature and the presence of catalysts can also change the rate of the reaction. Where do these other variables fit into the rate equation?



## 44.5 Effect of temperature on rate constants (p.53)

- ◆ These are included in the rate constant  $k$ , which increases with increasing temperature.





## 44.5 Effect of temperature on rate constants (p.53)

- ◆ The quantitative relationship between temperature and the rate constant is described by the **Arrhenius equation** (阿列紐斯方程式).

$$k = Ae^{\frac{-E_a}{RT}}, \text{ where}$$

- $k$  is the rate constant of the reaction;  
 $A$  is the collision factor;  
 $e$  is the base of the natural logarithm;  
 $E_a$  is the activation energy for the reaction in  $\text{J mol}^{-1}$ ;  
 $R$  is the gas constant (i.e.  $8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ ); and  
 $T$  is the temperature in Kelvin.



## 44.5 Effect of temperature on rate constants (p.53)

- ◆ A reaction that takes place fairly quickly at room temperature has an activation energy of about  $60 \text{ kJ mol}^{-1}$ .
- ◆ A temperature rise from  $298 \text{ K}$  to  $308 \text{ K}$  will cause the rate constant to increase by a factor of:

$$\frac{k_{308}}{k_{298}} = \frac{e^{-\frac{60\,000}{(8.31)(308)}}}{e^{-\frac{60\,000}{(8.31)(298)}}} = \frac{e^{-23.44}}{e^{-24.23}} = \frac{6.609 \times 10^{-11}}{3.000 \times 10^{-11}} = 2.2$$



## 44.5 Effect of temperature on rate constants (p.53)

Calculating activation energies graphically

- ◆ The Arrhenius equation can be expressed as

$$\ln k = -\frac{E_a}{RT} + \ln A \quad \text{or} \quad \log k = -\frac{E_a}{2.3RT} + \log A$$

▶ Notice that

$$\log x = \frac{\ln x}{\ln 10} = \frac{\ln x}{2.3}$$



## 44.5 Effect of temperature on rate constants (p.53)

$$\log k = -\frac{E_a}{2.3 R} \frac{1}{T} + \log A$$
$$y = m x + c$$

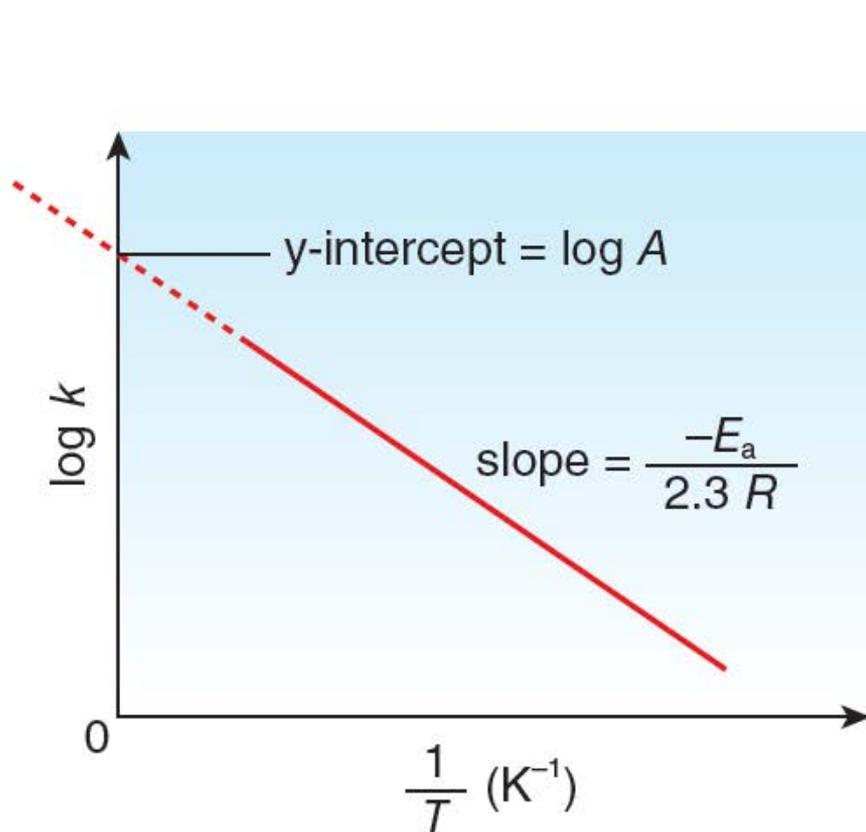
- ◆ A graph of  $\log k$  against  $\frac{1}{T}$  will be a straight line with a slope of  $-\frac{E_a}{2.3R}$  and a y-intercept of  $\log A$ .



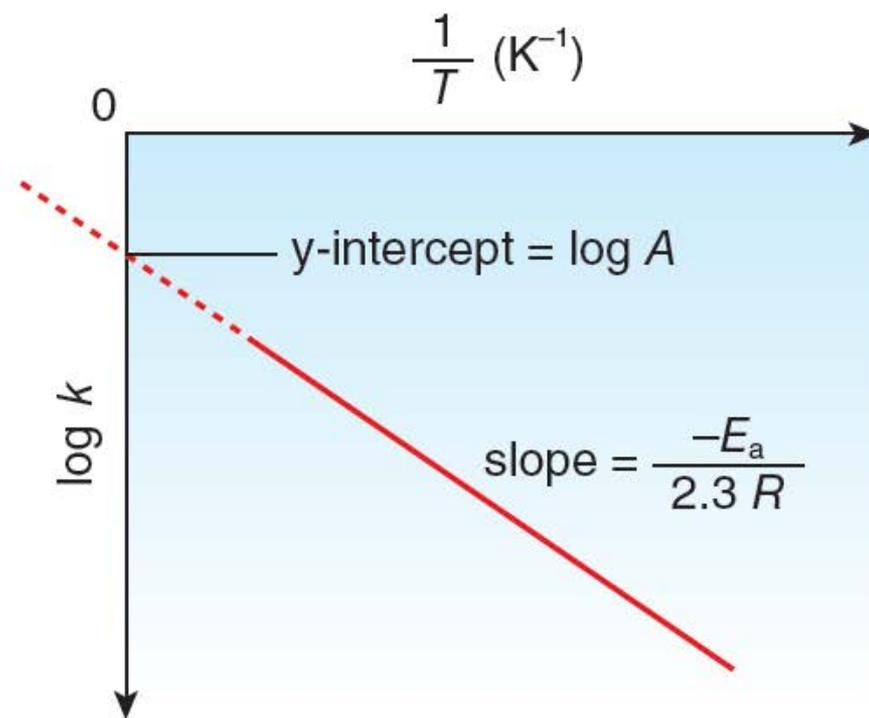
Determining the activation energy for a reaction [Ref.](#)



## 44.5 Effect of temperature on rate constants (p.53)



(a) If  $k$  is larger than 1,  $\log k$  is positive



(b) If  $k$  is smaller than 1,  $\log k$  is negative



## 44.5 Effect of temperature on rate constants (p.53)

Calculating activation energies non-graphically

$$\log k_1 = -\frac{E_a}{2.3RT_1} + \log A$$

$$\log k_2 = -\frac{E_a}{2.3RT_2} + \log A$$

$$\log k_1 - \log k_2 = -\frac{E_a}{2.3R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right)$$

$$\log \frac{k_2}{k_1} = \frac{E_a}{2.3R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right)$$



## 44.5 Effect of temperature on rate constants (p.53)

### Q (Example 44.1)

In a chemical kinetics experiment, the rate constants ( $k$ ) for the reaction of 1-bromopropane with hydroxide ion at various temperatures ( $T$ ) were determined.



The results were shown in the table below.

Temperature $T$ (K)	$k$ ( $\text{dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ )
298	$1.41 \times 10^{-4}$
308	$2.99 \times 10^{-4}$
318	$6.95 \times 10^{-4}$
328	$1.38 \times 10^{-3}$



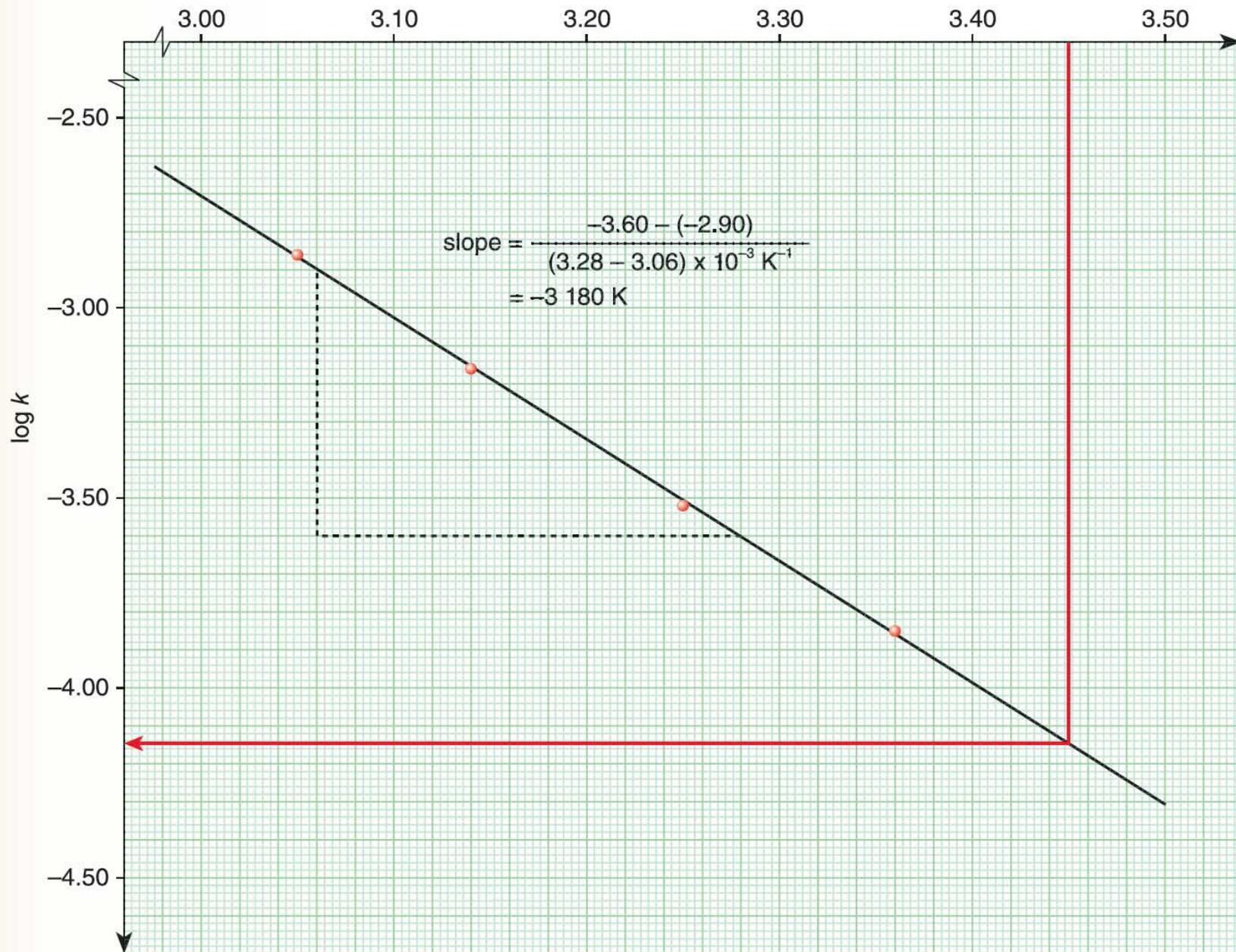
## 44.5 Effect of temperature on rate constants (p.53)

- Plot a graph of  $\log k$  on the vertical axis against  $\frac{1}{T}$  on the horizontal axis.
- Calculate the activation energy,  $E_a$ , for the reaction.
- Estimate the rate constant of the reaction at 290 K.  
(Gas constant  $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ )

A

a)

$\frac{1}{T} (\times 10^{-3} \text{ K}^{-1})$	$\log k$
3.36	-3.85
3.25	-3.52
3.14	-3.16
3.05	-2.86





## 44.5 Effect of temperature on rate constants (p.53)

b)

$$\text{Slope} = -\frac{E_a}{2.3 R} = \frac{-3.60 - (-2.90)}{(3.28 - 3.06) \times 10^{-3} \text{ K}^{-1}}$$
$$= -3\,180 \text{ K}$$

$$E_a = 2.3 \times 8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 3\,180 \text{ K}$$
$$= 60\,800 \text{ J mol}^{-1}$$
$$= 60.8 \text{ kJ mol}^{-1}$$

$\therefore$  the activation energy for the reaction is  $60.8 \text{ kJ mol}^{-1}$ .

c)

From the graph, at  $290 \text{ K}$  (i.e.  $\frac{1}{T} = 3.45 \times 10^{-3} \text{ K}^{-1}$ ),

$$\log k = -4.14$$

$$k = 7.24 \times 10^{-5} \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$$

$\therefore$  the rate constant at  $290 \text{ K}$  is  $7.24 \times 10^{-5} \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ .



## 44.5 Effect of temperature on rate constants (p.53)

### Q (Example 44.2)

The relationship between reaction rate constant  $k$  and absolute temperature  $T$  can be represented by the Arrhenius equation:

$$k = Ae^{\frac{-E_a}{RT}}, \text{ where } A \text{ is the collision factor;}$$

$E_a$  is the activation energy for the reaction in  $\text{J mol}^{-1}$ ;  
 $R$  is the gas constant.

Given that the rate constant of a reaction is  $2.60 \times 10^5 \text{ cm}^3 \text{ mol}^{-1} \text{ s}^{-1}$  at 210 K and  $A$  is  $4.80 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ , calculate its  $E_a$ .  
(Gas constant  $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ ).



## 44.5 Effect of temperature on rate constants (p.53)

A

$$\log k = -\frac{E_a}{2.3 RT} + \log A$$

$$\frac{E_a}{2.3 RT} = \log \left( \frac{A}{k} \right)$$

$$E_a = 2.3 RT \log \left( \frac{A}{k} \right)$$

$$= 2.3 \times 8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 210 \text{ K} \times \log \left( \frac{4.80 \times 10^9}{2.60 \times 10^5} \right)$$

$$= 17\,100 \text{ J mol}^{-1}$$

$$= 17.1 \text{ kJ mol}^{-1}$$

$\therefore$  the  $E_a$  for the reaction is  $17.1 \text{ kJ mol}^{-1}$ .



## 44.5 Effect of temperature on rate constants (p.53)

### Q (Example 44.3)

The rate constant of a certain reaction doubles when the temperature is increased from 308 K to 318 K.

Calculate the activation energy for the reaction.

(Gas constant  $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ )

### A

$$k = Ae^{\frac{-E_a}{RT}}$$

$$\log \frac{k_{318}}{k_{308}} = \log 2 = -\frac{E_a}{2.3R} \left( \frac{1}{308} - \frac{1}{318} \right)$$

$$E_a = 56.4 \text{ kJ mol}^{-1}$$

$\therefore$  the activation energy for the reaction is  $56.4 \text{ kJ mol}^{-1}$



## 44.5 Effect of temperature on rate constants (p.53)

### Practice 44.3

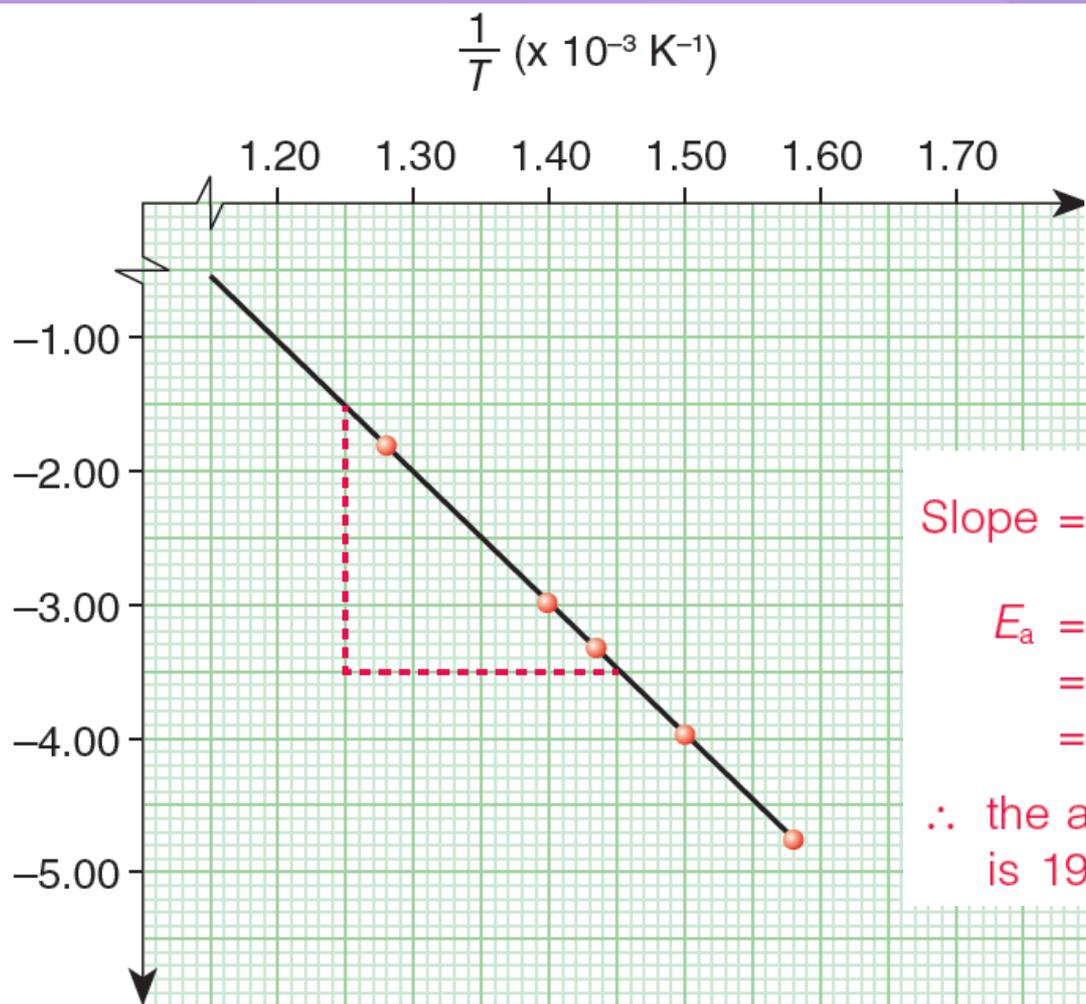
1 In a chemical kinetics experiment, the rate constants ( $k$ ) of a reaction at various temperatures ( $T$ ) were determined.

The graph below shows the plot of  $\log k$  against  $\frac{1}{T}$ . Calculate the activation energy for this reaction.

(Gas constant  $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ )



## 44.5 Effect of temperature on rate constants (p.53)



$$\text{slope} = \frac{-3.50 - (-1.50)}{(1.45 - 1.25) \times 10^{-3} \text{ K}^{-1}} = -10\,000 \text{ K}$$

$$\text{Slope} = -\frac{E_a}{2.3R} = -10\,000 \text{ K}$$

$$\begin{aligned} E_a &= 2.3 \times 8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 10\,000 \text{ K} \\ &= 191\,000 \text{ J mol}^{-1} \\ &= 191 \text{ kJ mol}^{-1} \end{aligned}$$

$\therefore$  the activation energy for the reaction is  $191 \text{ kJ mol}^{-1}$ .



## 44.5 Effect of temperature on rate constants (p.53)

2 The rate constant of a certain reaction quadruples when the temperature is increased from 298 K to 328 K.

Calculate the activation energy for the reaction.

(Gas constant  $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ )

$$k = Ae^{-\frac{E_a}{RT}}$$

$$\log \frac{k_{328}}{k_{298}} = \log 4 = \frac{E_a}{2.3R} \left( \frac{1}{298} - \frac{1}{328} \right)$$

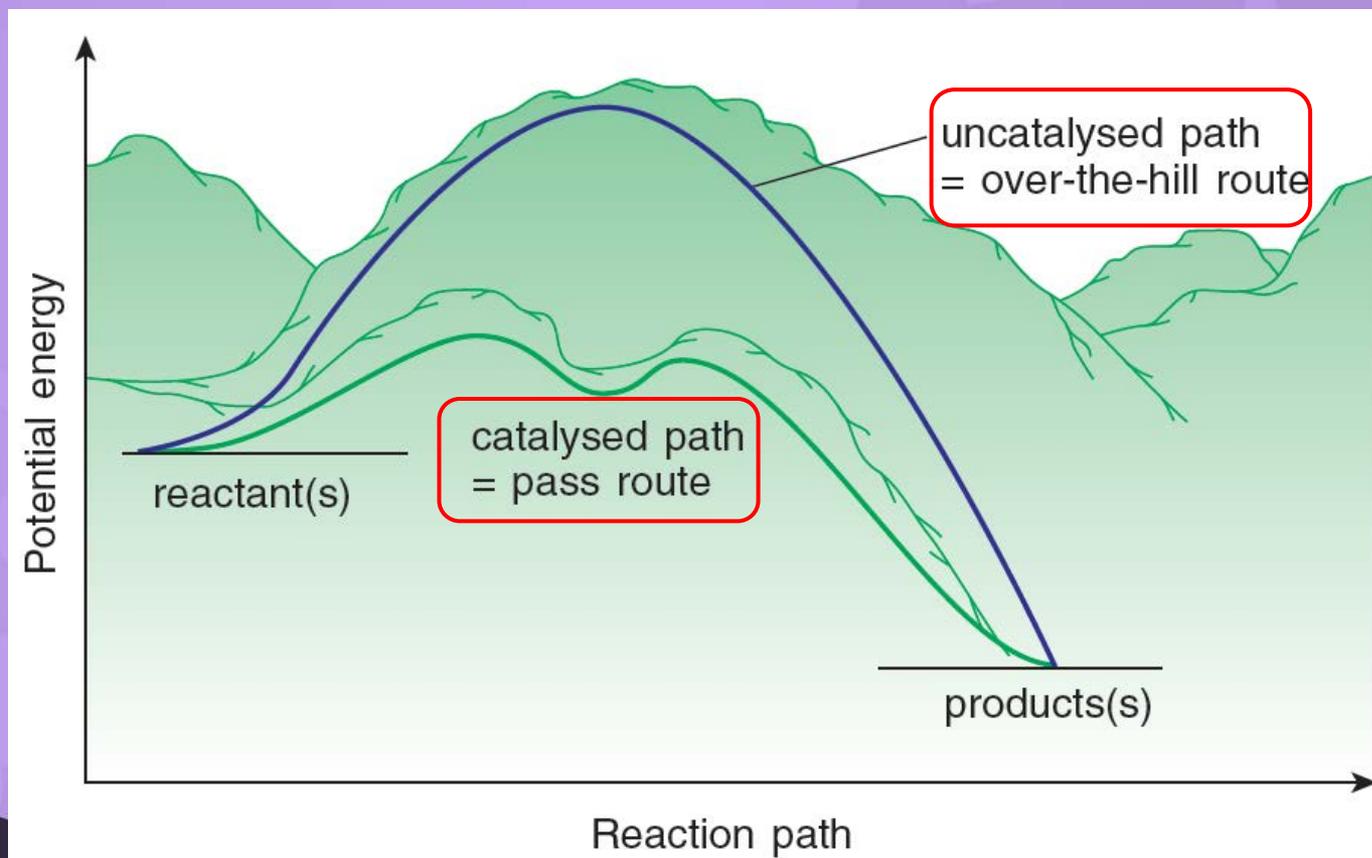
$$E_a = 37.5 \text{ kJ mol}^{-1}$$

$\therefore$  the activation energy for the reaction is  $37.5 \text{ kJ mol}^{-1}$ .



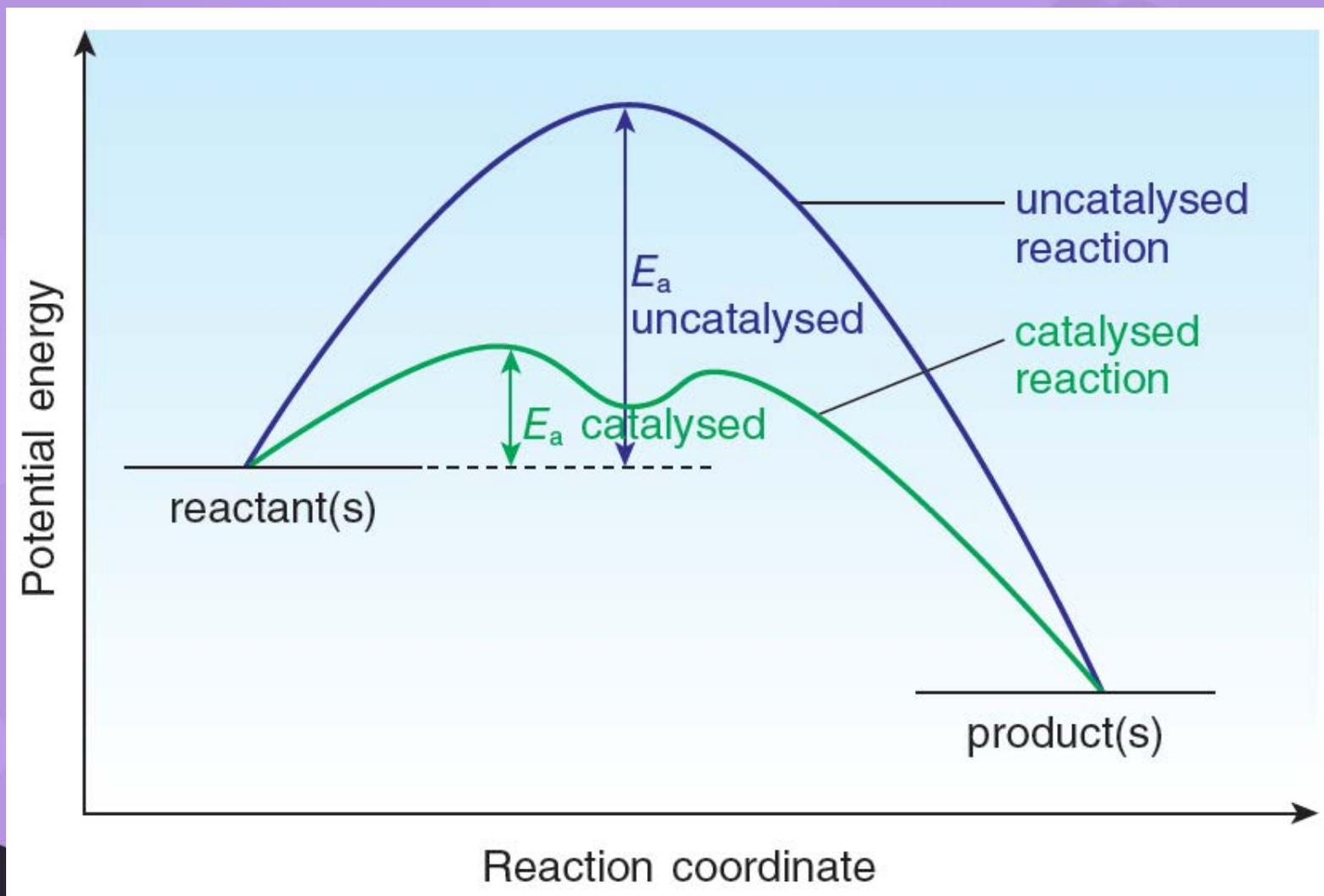
## 44.6 What does a catalyst do? (p.60)

- ◆ A catalyst increases the rate of a chemical reaction by providing an alternative reaction pathway of lower activation energy.





## 44.6 What does a catalyst do? (p.60)





## 44.6 What does a catalyst do? (p.60)

- ◆ Consider a reaction that takes place at 323 K. The activation energy of the reaction is 50 kJ mol<sup>-1</sup>. A catalyst provides an alternative reaction pathway of 40 kJ mol<sup>-1</sup>.

$$k_{\text{catalysed}} = Ae^{-\frac{40\,000}{(8.31)(323)}}$$

$$k_{\text{uncatalysed}} = Ae^{-\frac{50\,000}{(8.31)(323)}}$$

$$\frac{k_{\text{catalysed}}}{k_{\text{uncatalysed}}} = \frac{e^{-\frac{40\,000}{(8.31)(323)}}}{e^{-\frac{50\,000}{(8.31)(323)}}}$$
$$= 41.6$$

- ◆ A decrease of 10 kJ mol<sup>-1</sup> in the activation energy significantly increases the rate constant and the rate of the reaction to 41.6 times.



## 44.6 What does a catalyst do? (p.60)

### Characteristics of catalysts

- ◆ Only small amounts of a catalyst are usually needed.
- ◆ A specific catalyst is often needed for one particular reaction. This is particularly true for enzymes.
- ◆ A catalyst does not affect the enthalpy change of a chemical reaction.
- ◆ A catalyst lowers the activation energy of both the forward and backward reactions of a reversible reaction by the same amount. Consequently the presence of a catalyst increases the rate of the forward and backward reactions equally.



## 44.6 What does a catalyst do? (p.60)

- ◆ A catalyst does not change the equilibrium constant and the position of equilibrium. It only increases the rate at which equilibrium is attained and hence decreases the time required to attain equilibrium.
- ◆ In other words, the presence of a catalyst does not increase the yield of products but increases the rate of their production.



## 44.6 What does a catalyst do? (p.60)

### Catalyst poisoning

- ◆ Many catalysts are poisoned by trace impurities. For example, catalysts for the hydrogenation process can be poisoned by traces of sulphur impurities.
- ◆ Another example of catalyst poisoning is the use of leaded petrol in cars with a catalytic converter that would normally take only unleaded petrol. The lead in the petrol coats the surface of the catalyst in the converter and makes it ineffective as a catalyst.



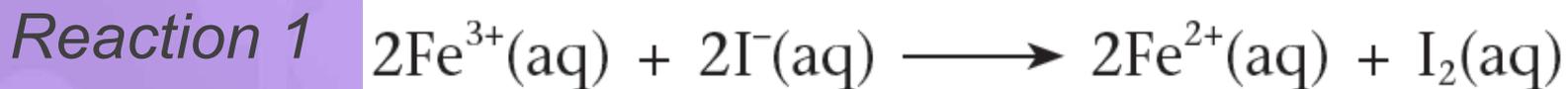
## 44.6 What does a catalyst do? (p.60)

### Catalysis by transition metal

- ◆ Peroxodisulphate ion ( $\text{S}_2\text{O}_8^{2-}$ ) oxidises iodide ion to iodine according to the following equation:



- ◆ The reaction is quite slow. Both ions are negatively charged, so they are unlikely to make effective collisions with each other.
- ◆  $\text{Fe}^{3+}(\text{aq})$  ion catalyses this reaction. The catalysis involves two redox reactions.





## 44.6 What does a catalyst do? (p.60)

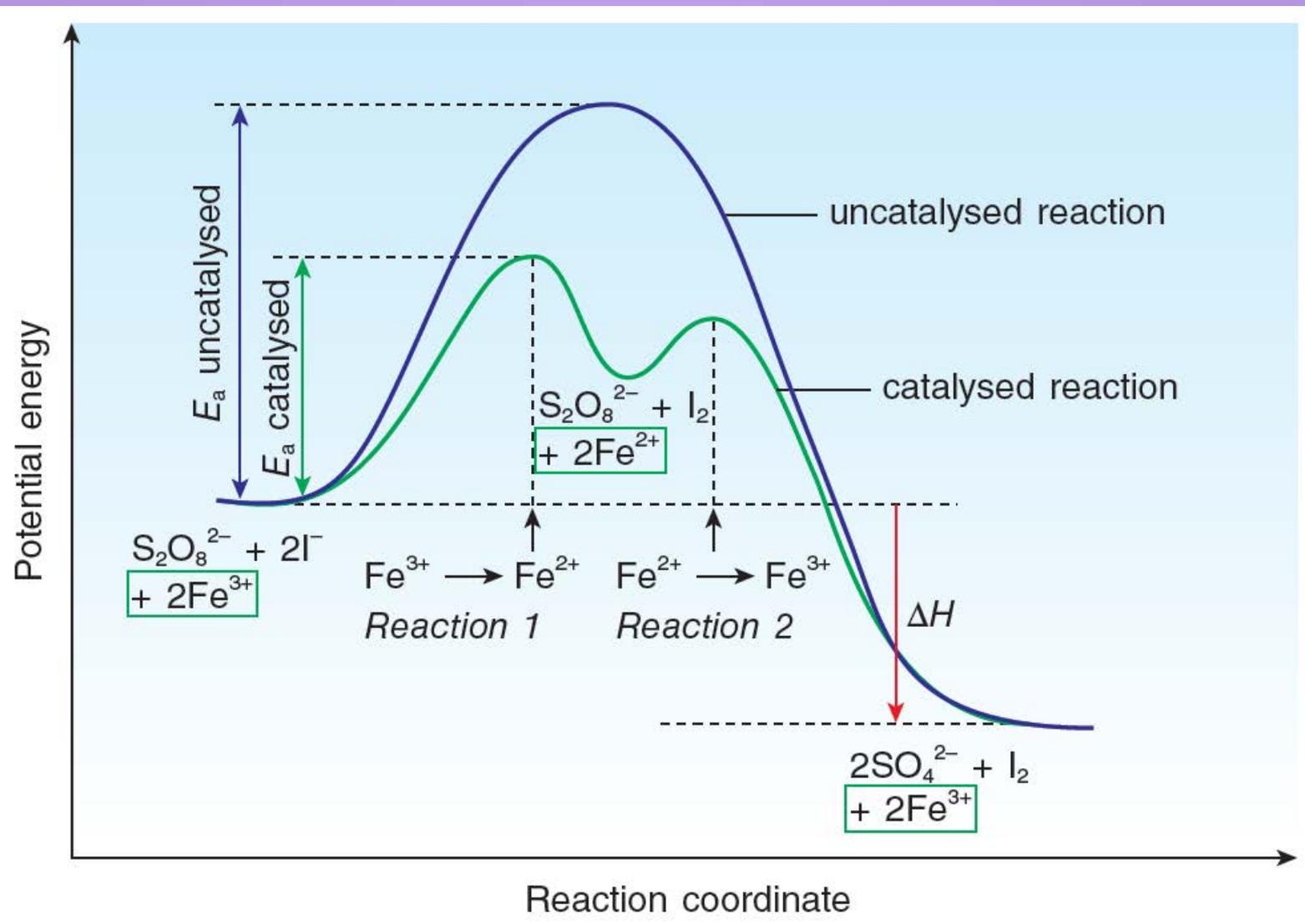
- ◆ In both *Reactions 1* and *2*, positively charged ions of iron react with negatively charged ions. The reaction of ions with unlike charges are more likely to occur than the direct reaction between  $\text{S}_2\text{O}_8^{2-}(\text{aq})$  and  $\text{I}^-(\text{aq})$  ions.
- ◆ Although the iron(III) ion is used up in *Reaction 1*, it is regenerated in *Reaction 2*. Overall the iron(III) ion is not consumed.



Catalysing the reaction  
between peroxodisulphate ion  
and iodide ion



## 44.6 What does a catalyst do? (p.60)





## 44.6 What does a catalyst do? (p.60)

### Practice 44.4

Sulphur dioxide is produced when fossil fuels containing sulphur are burnt. When sulphur dioxide escapes into the atmosphere, it is oxidised to sulphur trioxide.



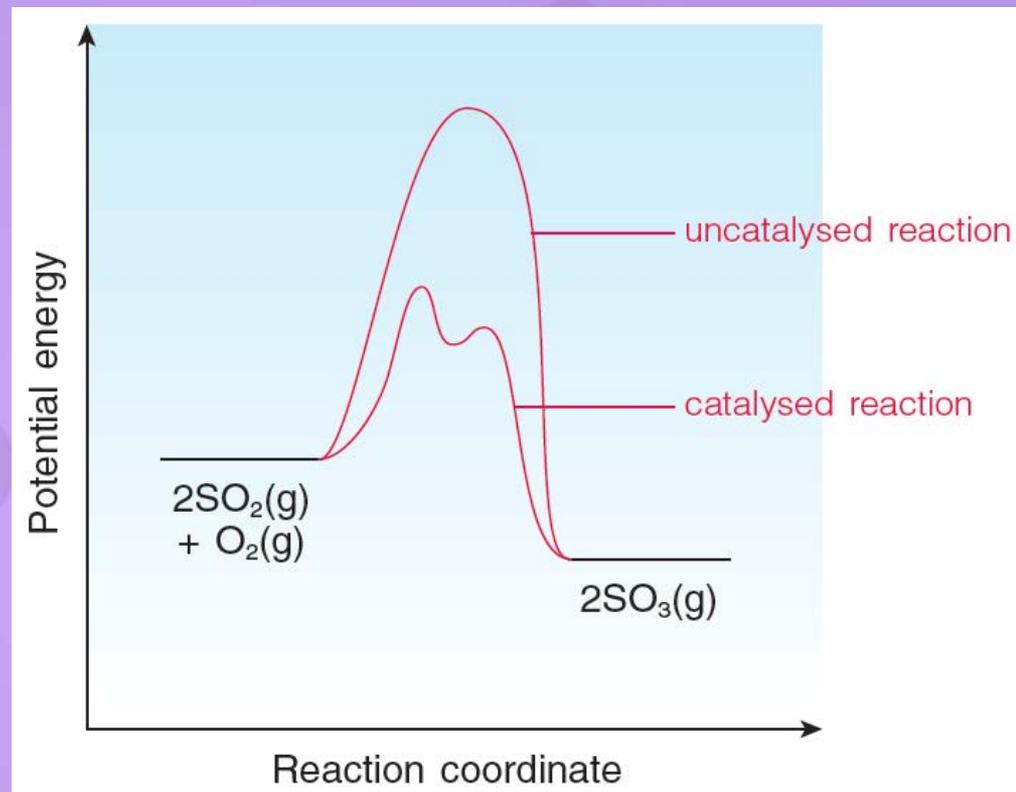
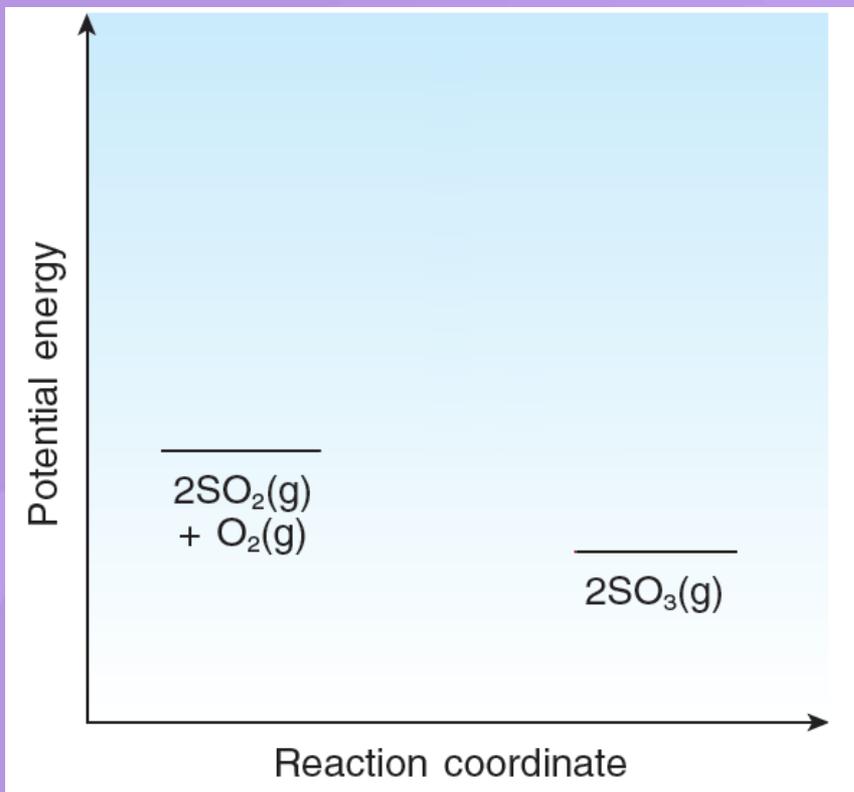
Nitrogen dioxide present in the atmosphere can catalyse the oxidation of sulphur dioxide. The following two reactions occur in the catalytic process:



Sketch the potential energy profiles for the uncatalysed and catalysed conversions.



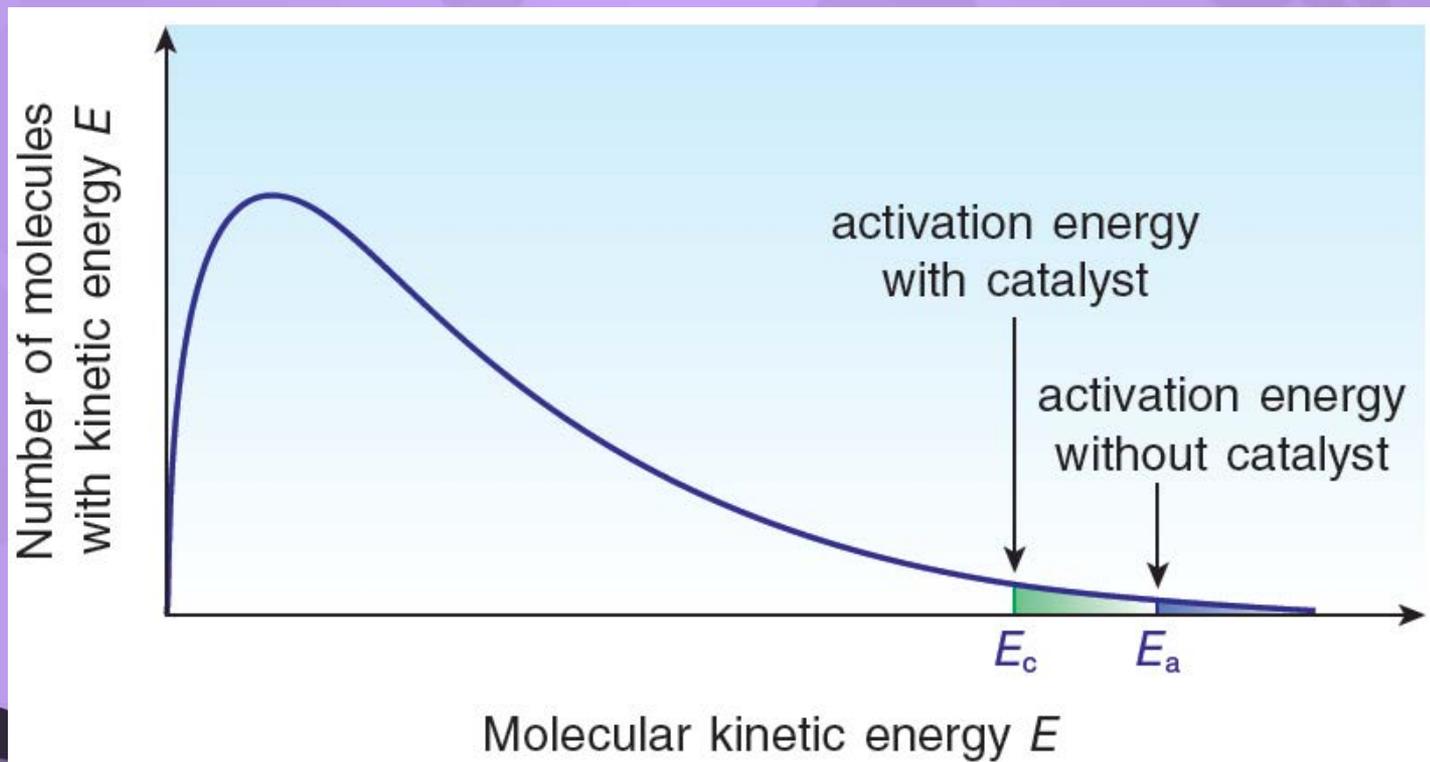
## 44.6 What does a catalyst do? (p.60)





## 44.7 The Maxwell-Boltzmann distribution curve and catalyst (p.65)

- ◆ A catalyst provides an alternative reaction pathway of lower activation energy ( $E_c$ ), as compared to the the case without catalyst ( $E_a$ ).





## 44.7 The Maxwell-Boltzmann distribution curve and catalyst (p.65)

- ◆ A greater fraction of molecules now has kinetic energy equal to or greater than the lower activation energy,  $E_c$ . Hence more molecules will collide and react to form products. The result is an increase in the rate of reaction.

key:



area representing fraction of molecules with kinetic energy equal to or greater than activation energy without a catalyst ( $E_a$ )



area representing fraction of molecules with kinetic energy equal to or greater than activation energy with a catalyst ( $E_c$ )



## 44.7 The Maxwell-Boltzmann distribution curve and catalyst (p.65)

### Practice 44.5

The decomposition of ammonia is represented by the equation below.



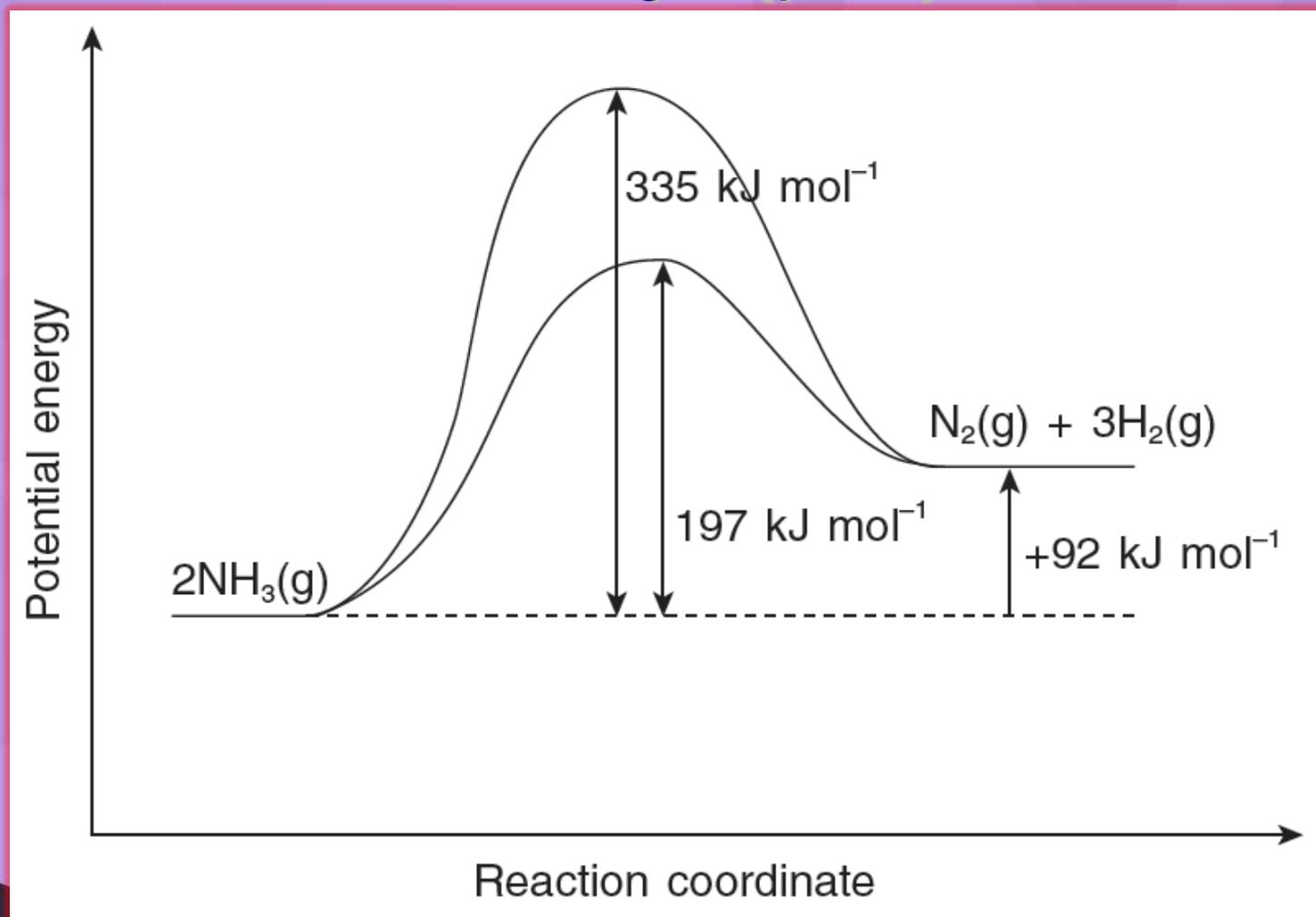
The activation energy for the uncatalysed reaction is  $335 \text{ kJ mol}^{-1}$ .

The activation energy for the reaction when osmium is used as a catalyst is  $197 \text{ kJ mol}^{-1}$ .

a) Sketch the potential energy profiles for the uncatalysed and catalysed reactions.



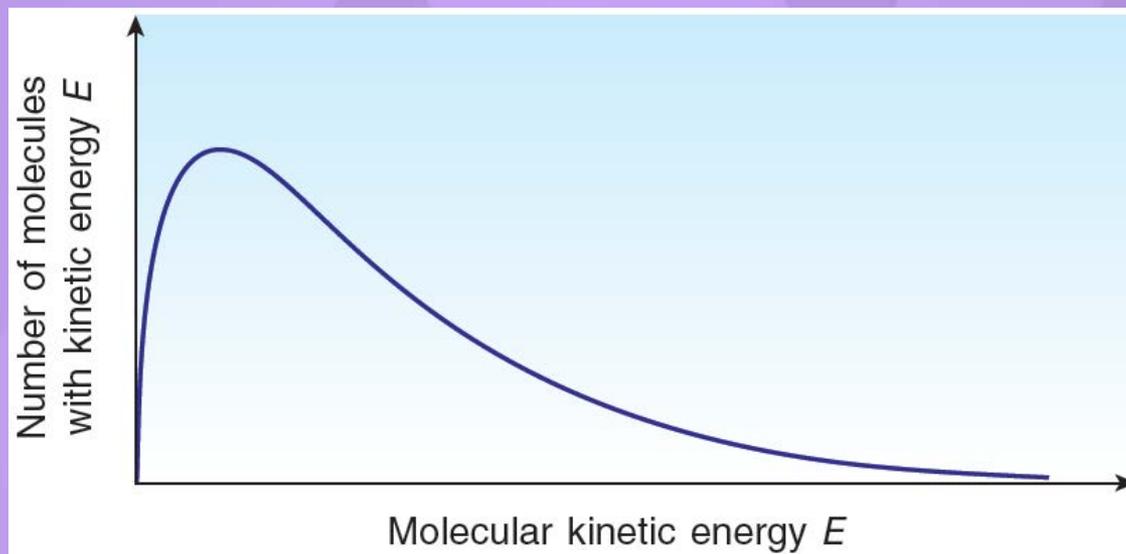
## 44.7 The Maxwell-Boltzmann distribution curve and catalyst (p.65)





## 44.7 The Maxwell-Boltzmann distribution curve and catalyst (p.65)

- b) When tungsten is used as a catalyst, the activation energy is  $163 \text{ kJ mol}^{-1}$ . The diagram below shows the Maxwell-Boltzmann distribution curve of molecular kinetic energies of a gas at a certain temperature.



With reference to the curve, explain which catalyst, osmium or tungsten, will cause ammonia to decompose at a higher rate.



## 44.7 The Maxwell-Boltzmann distribution curve and catalyst (p.65)

When tungsten is used, the activation energy is lower. Compared with using osmium, a greater fraction of molecules have kinetic energy equal to or greater than this lower activation energy. Hence more molecules have sufficient energy to collide and decompose to form products.



## 44.8 Catalysts in industry (p.66)

The two major economic advantages of the use of catalysts are:

- ◆ they increase the rate of chemical reactions, meaning that more of the desired products can be made in a given time period;
- ◆ reactions can take place at lower temperatures resulting in a reduction in the energy costs to the manufacturer.

e.g. finely divided iron (Haber process for  $\text{NH}_3$ );  
vanadium(V) oxide (Contact process for  $\text{H}_2\text{SO}_4$ ).



## 44.9 Enzymes (p.68)

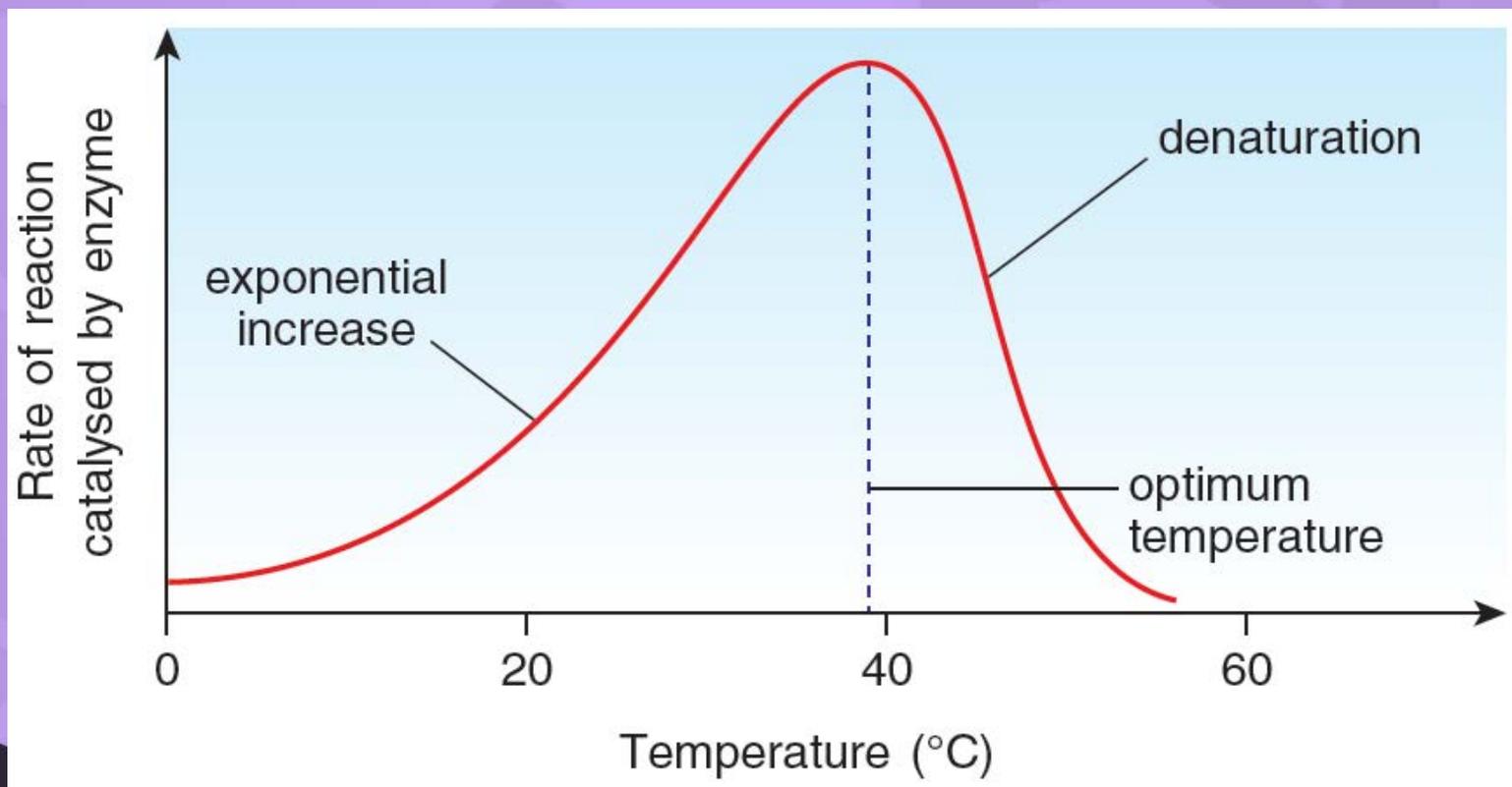
Enzymes are biological catalysts. They are large protein molecules which enable the biochemical reactions within the body of living organisms to take place very quickly at relatively low temperatures. Enzymes

- ◆ speed up a reaction without themselves being used up;
- ◆ provide an alternative reaction pathway with a lower activation energy.
- ◆ are very specific — each enzyme only catalyses one particular reaction.
- ◆ are sensitive to pH — many work best at a particular pH (usually around 7) and become inactive (denatured) at pH values that are further from the optimum.



## 44.9 Enzymes (p.68)

- are sensitive to temperature — many enzymes work best at temperatures close to body temperature and most are denatured at 60–70 °C.

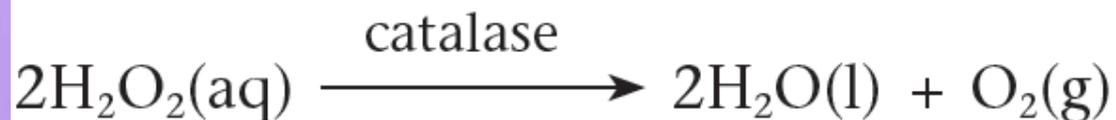




## 44.9 Enzymes (p.68)

### Enzymes in biological processes

- ◆ Amylase found in saliva starts to break down starchy food into maltose in the mouth.
- ◆ Protease breaks down proteins into amino acids in the small intestine.
- ◆ Catalase found in the liver catalyses the reaction below to prevent a build-up of hydrogen peroxide, which is potentially harmful to cells.

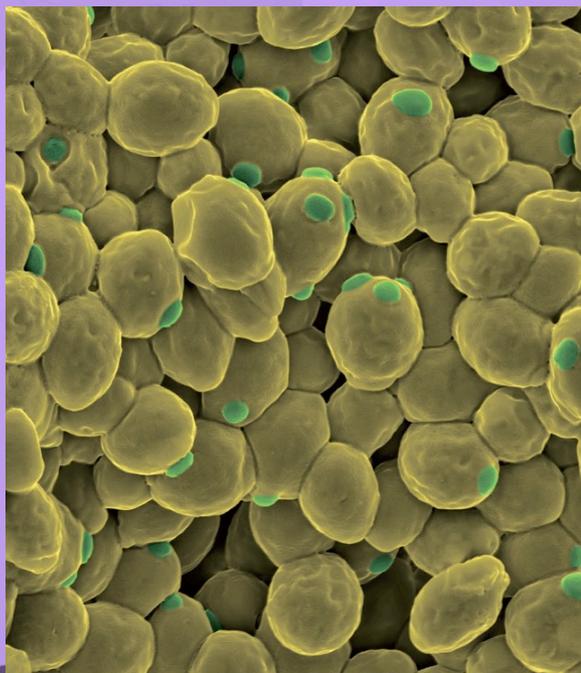


## 44.9 Enzymes (p.68)

### Preparing ethanol by fermentation

- ◆ In the U.S.A., most ethanol is produced by fermentation of carbohydrates (from corn) with enzymes of yeast. Ethanol produced by this method is mainly used as a fuel, with smaller amounts used in alcoholic beverages and the chemical industry.

**Brewer's yeast  
under scanning  
electron microscope**





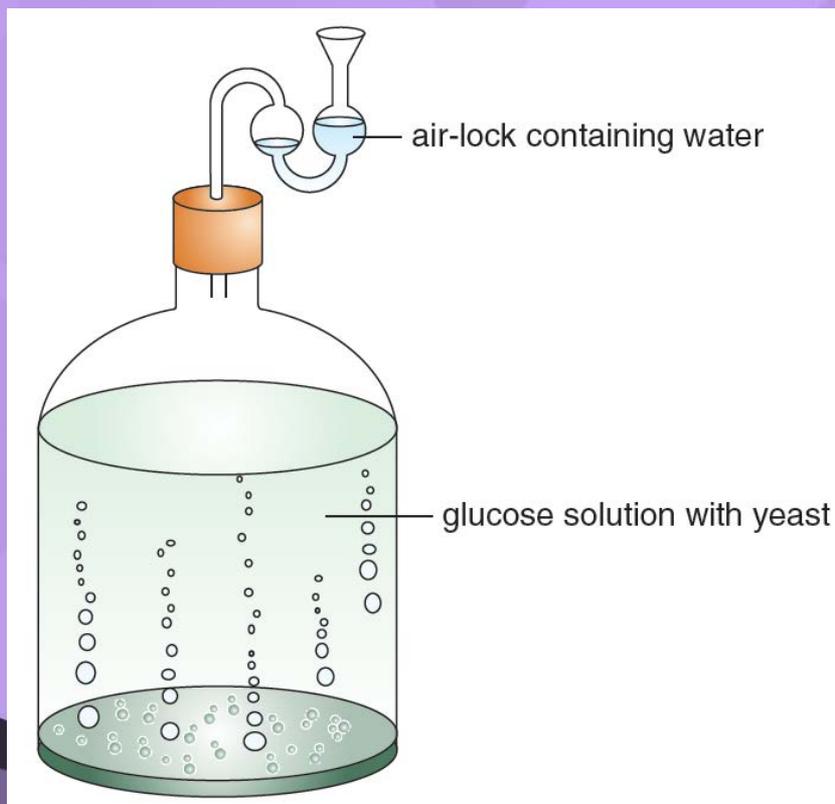


## 44.9 Enzymes (p.68)

- ◆ The optimum temperature for the fermentation reaction with yeast is in the range 25–38 °C. At lower temperatures the rate of reaction is too low, and at higher temperatures the enzymes are denatured.
- ◆ The conditions required for successful fermentation are:
  - yeast;
  - water;
  - yeast nutrients (ammonium phosphate is often used);
  - warmth (a temperature of 25–38 °C);
  - absence of air (with oxygen present, the ethanol can be partially oxidised to ethanoic acid, or completely oxidised to carbon dioxide).

 44.9 Enzymes (p.68)

- ◆ The reaction stops when the ethanol concentration reaches about 15%. Above this concentration, the yeast cells die. After filtering off the dead yeast cells, the solution is fractionally distilled. The distillate consists of 95% ethanol and 5% water.



**A laboratory  
fermentation vessel**



## 44.9 Enzymes (p.68)

Enzymes are becoming more widely used as industrial catalysts because:

- ◆ they usually work at best at quite low temperatures and thus save money on heating;
- ◆ they operate in mild conditions (in water and usually close to pH 7) that do not harm fabrics or food;
- ◆ they can be controlled easily as they stop working when the temperature or pH is changed;
- ◆ they often allow a reaction which forms pure products (with no side reactions) to take place, and this often eliminates the need for complex separation techniques;



## 44.9 Enzymes (p.68)

- ♦ they are biodegradable and thus disposal of waste enzymes does not cause environmental problems.



## 44.9 Enzymes (p.68)

### Practice 44.6

Ethanol is formed by the enzyme-catalysed fermentation of glucose.

a) A student carried out this fermentation reaction in a beaker using an aqueous solution of glucose at 25 °C in the presence of yeast.

i) Write the equation for the fermentation reaction that takes place.



ii) The ethanol obtained had a sour odour. The student suspected the product was contaminated by ethanoic acid. How did this contamination occur? **Air got in. / Oxidation occurred.**

b) One benefit of using an enzyme to catalyse a reaction is that there is a large increase in the rate of reaction without the enzyme being used up. State TWO other benefits, other than cost, of using enzymes to catalyse reactions. **Any two of the following:**

- Lower temperature can be used. / Carbon dioxide emissions from burning of fossil fuels can be reduced.
- Enzymes are very specific.
- The use of toxic solvents / toxic catalysts can be reduced.



## Key terms (p.73)

activation energy	活化能	Maxwell-Boltzmann distribution	麥克斯韋-波爾茲曼分佈
potential energy profile	勢能能線圖	most probable kinetic energy	最大機率動能
transition state	過渡態	Arrhenius equation	阿列紐斯方程式
activated complex	活化複體		

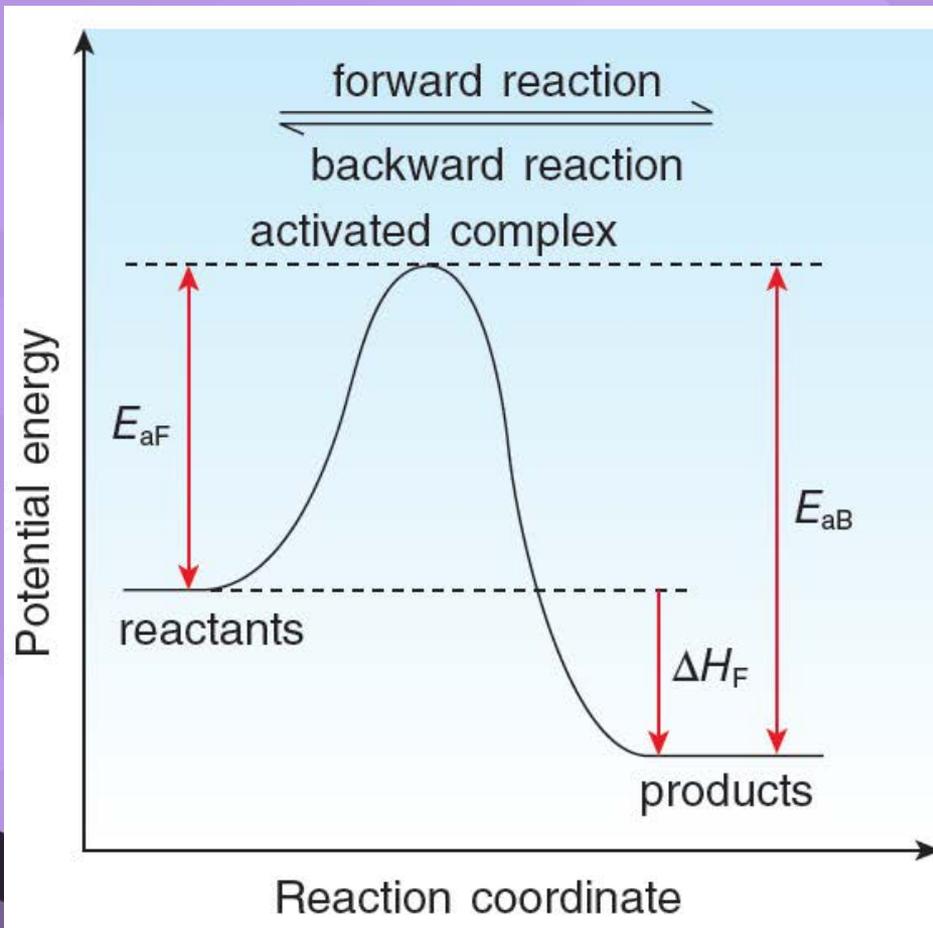


## Summary (p.74)

1 Activation energy is the minimum kinetic energy that colliding particles must have for a collision to result in a chemical reaction.

 Summary (p.74)

2 The following diagram shows the potential energy profile for a reversible reaction whose forward reaction is exothermic.



key:

$E_{aF}$  = activation energy for the forward reaction

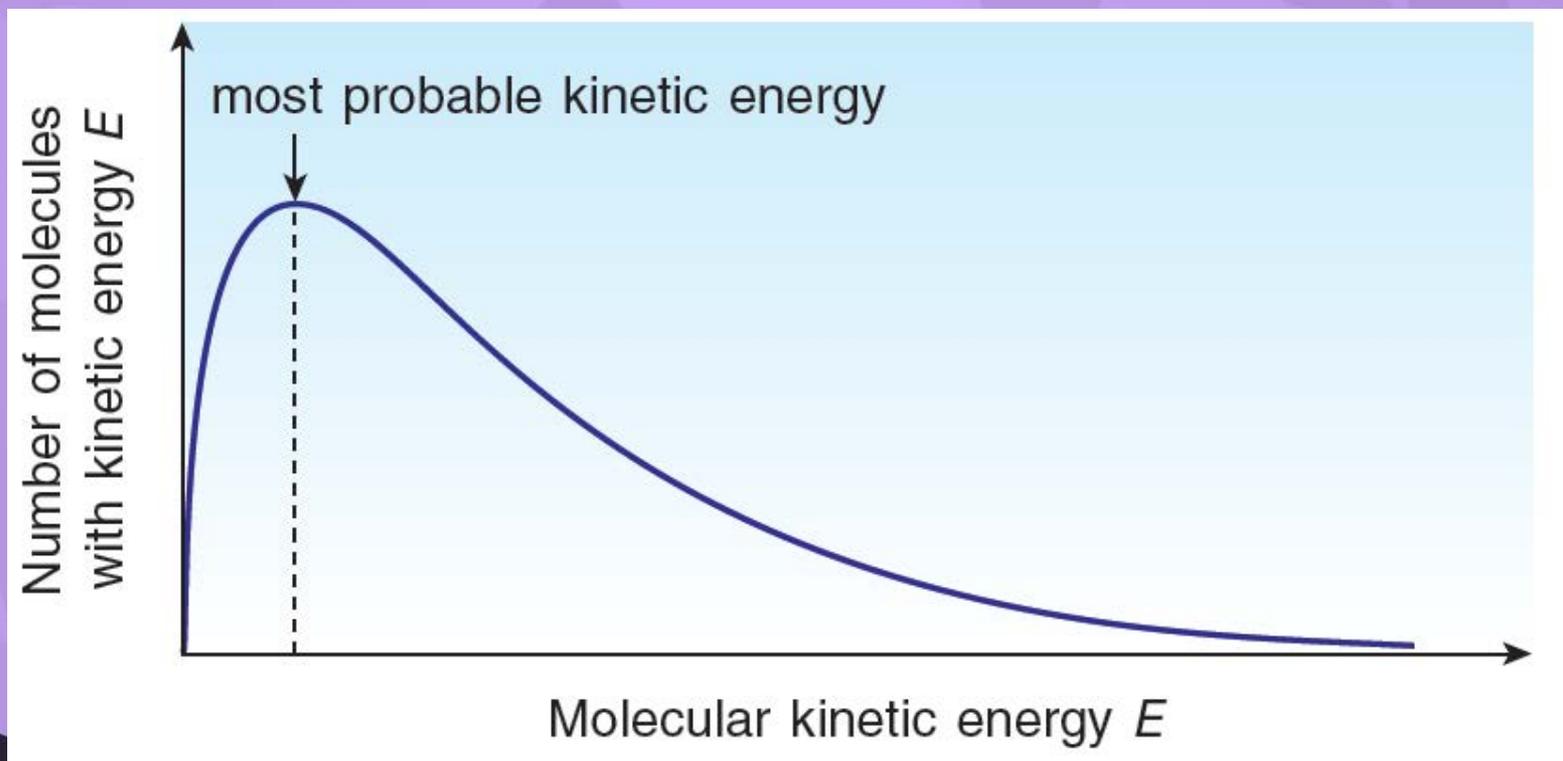
$E_{aB}$  = activation energy for the backward reaction

$\Delta H_F$  = enthalpy change of the forward reaction



## Summary (p.74)

- 3 The distribution of molecular kinetic energy among molecules in a gas at a certain temperature is called the Maxwell-Boltzmann distribution.





## Summary (p.74)

4 In summary, a temperature rise

- increases the average kinetic energy of the molecules, causing the collision frequency to increase, and hence effective collisions among molecules occur more frequently (accounts for a small increase in the rate of reaction).
- increases the fraction of molecules with kinetic energy equal to or greater than the activation energy, and hence a higher percentage of molecules with sufficient energy to collide and react (accounts for a rapid increase in the rate of reaction).



## Summary (p.74)

5 a) The Arrhenius equation is

$$k = Ae^{\frac{-E_a}{RT}}, \text{ where}$$

- $k$  is the rate constant of the reaction;
- $A$  is the collision factor;
- $e$  is the base of the natural logarithm;
- $E_a$  is the activation energy for the reaction in  $\text{J mol}^{-1}$ ;
- $R$  is the gas constant (i.e.  $8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ ); and
- $T$  is the temperature in Kelvin.



## Summary (p.74)

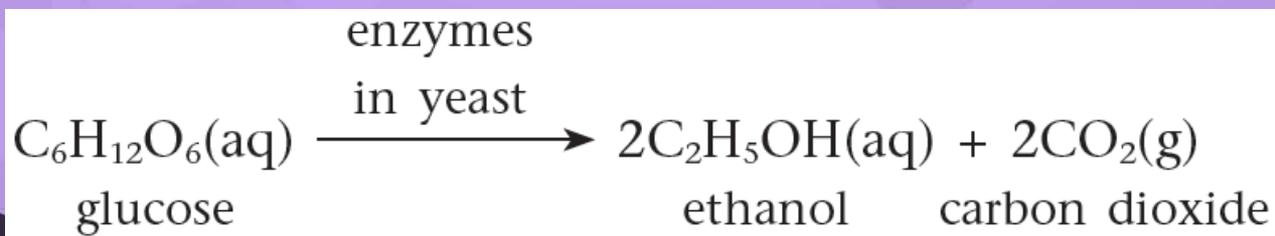
b) A graph of  $\log k$  against  $\frac{1}{T}$  is a straight line with a slope of  $-\frac{E_a}{2.3R}$  and a y-intercept of  $\log A$ .

The value of the activation energy,  $E_a$ , can be calculated from the slope.



## Summary (p.74)

- 6 A catalyst increases the rate of a chemical reaction by providing an alternative reaction pathway of lower activation energy.
- 7 Enzymes are biological catalysts. They are large protein molecules which enable the biochemical reactions that happen in living things to take place very quickly at relatively low temperatures.
- 8 Yeast cells contain enzymes that catalyse the conversion of glucose to ethanol and carbon dioxide.



## Unit Exercise (p. 76)

**Note:** Questions are rated according to ascending level of difficulty (from 1 to 5):

 question targeted at level 3 and above;

 question targeted at level 4 and above;

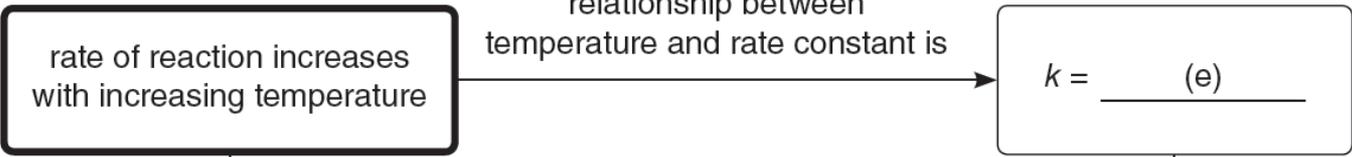
 question targeted at level 5.

**' \* '** indicates 1 mark is given for effective communication.

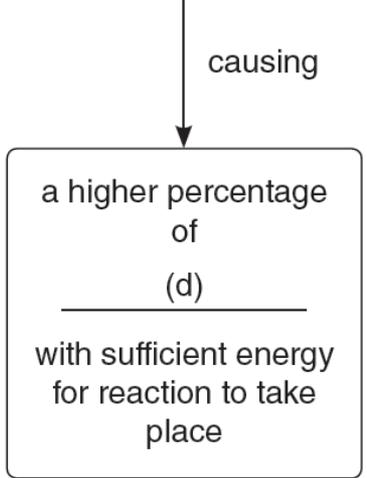
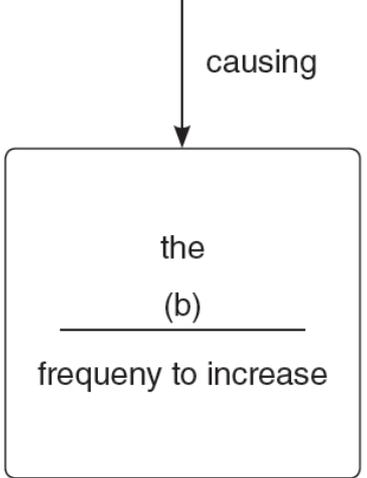
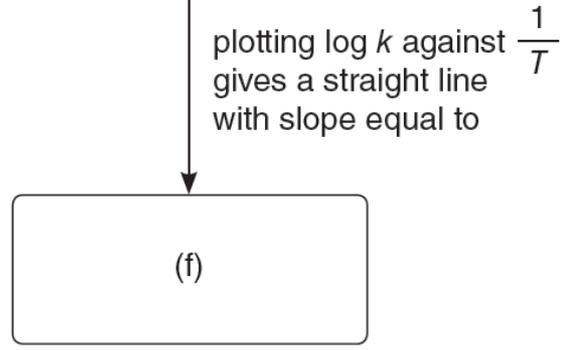
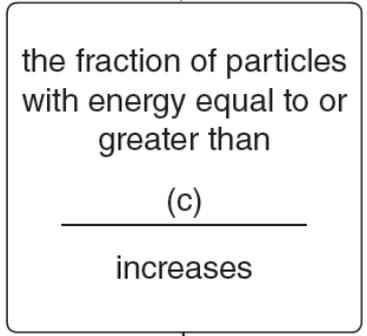
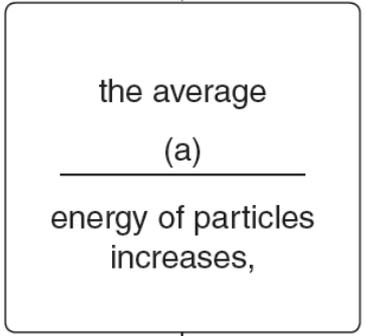
 Unit Exercise (p. 76)

**PART I KNOWLEDGE AND UNDERSTANDING**

1 Complete the following concept map.



because



- a) kinetic
- b) collision
- c) activation energy
- d) collisions
- e)  $Ae^{\frac{-E_a}{RT}}$
- f)  $-\frac{E_a}{2.3R}$

 Unit Exercise (p. 76)**PART II MULTIPLE CHOICE QUESTIONS**

2 How do the kinetic and potential energies change as reactant molecules approach each other to form an activated complex?

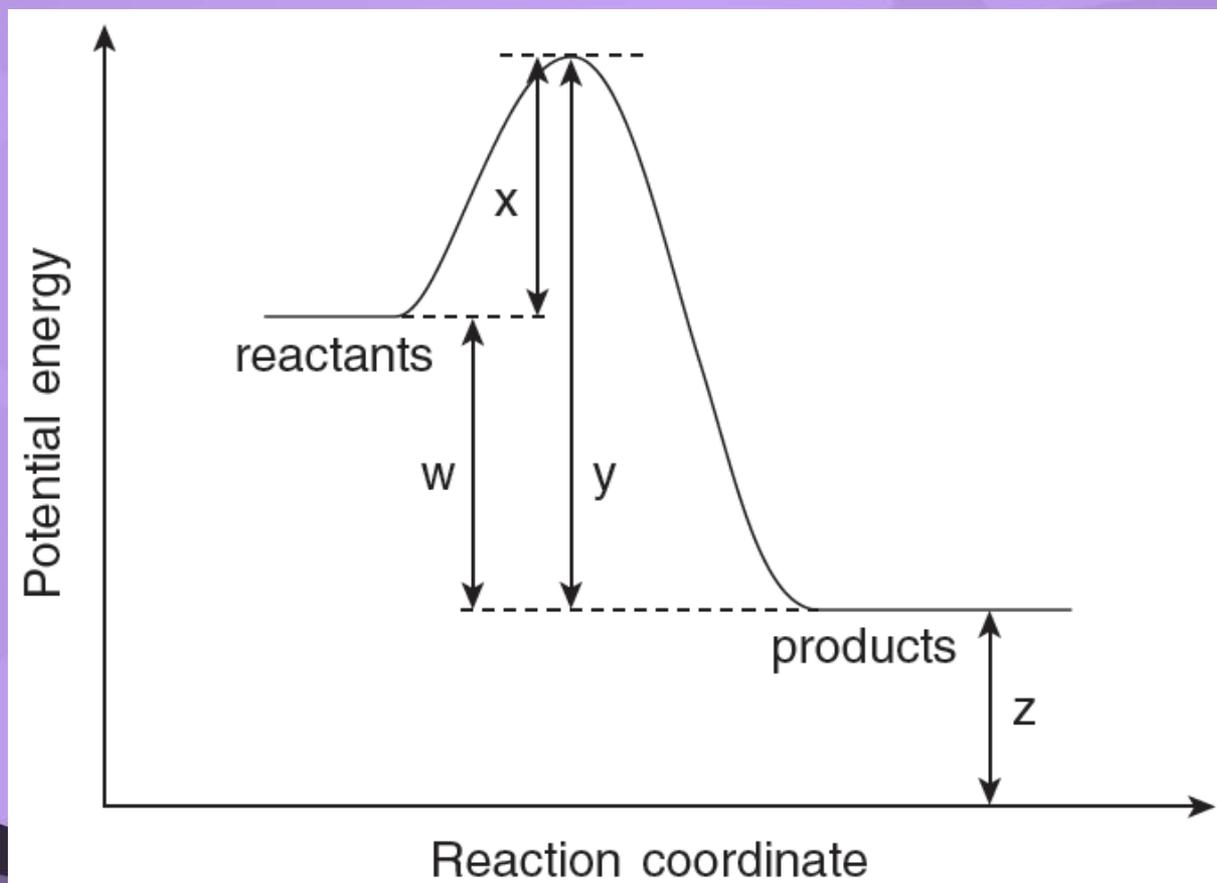
	<u>Kinetic energy</u>	<u>Potential energy</u>
A	increases	decreases
B	decreases	increases
C	decreases	remains constant
D	remains constant	remains constant

**Answer: B**

 Unit Exercise (p. 76)

- 3 The potential energy profile for a reversible reaction is shown below. Which of the arrows represents the activation energy for the backward reaction?

- A w
- B x
- C y
- D z



Answer: C



## Unit Exercise (p. 76)

4 Which of the following is true for an endothermic reversible reaction?



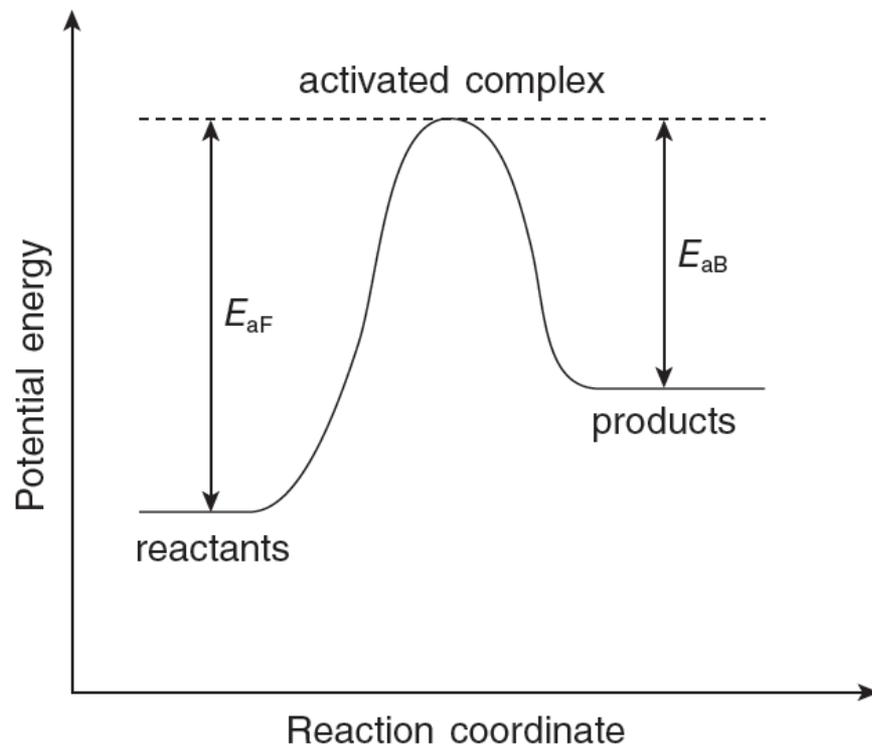
- A activation energy for forward reaction  $>$  activation energy for backward reaction
- B activation energy for forward reaction  $<$  activation energy for backward reaction
- C activation energy for forward reaction  $=$  activation energy for backward reaction
- D activation energy for forward reaction  $+$  activation energy for backward reaction  $= 0$

**Answer: A**

 Unit Exercise (p. 76)

## Explanation:

A The potential energy profile for an endothermic reversible reaction is shown below.



key:

$E_{aF}$  = activation energy for forward reaction

$E_{aB}$  = activation energy for backward reaction

$\therefore$  activation energy for forward reaction > activation energy for backward reaction



## Unit Exercise (p. 76)

5 What would happen to the rate constant ( $k$ ) and activation energy ( $E_a$ ) of a  reaction when the temperature is increased?

	<u>Rate constant (<math>k</math>)</u>	<u>Activation energy (<math>E_a</math>)</u>
A	increases	remains unchanged
B	decreases	remains unchanged
C	remains unchanged	increases
D	remains unchanged	decreases

Answer: A

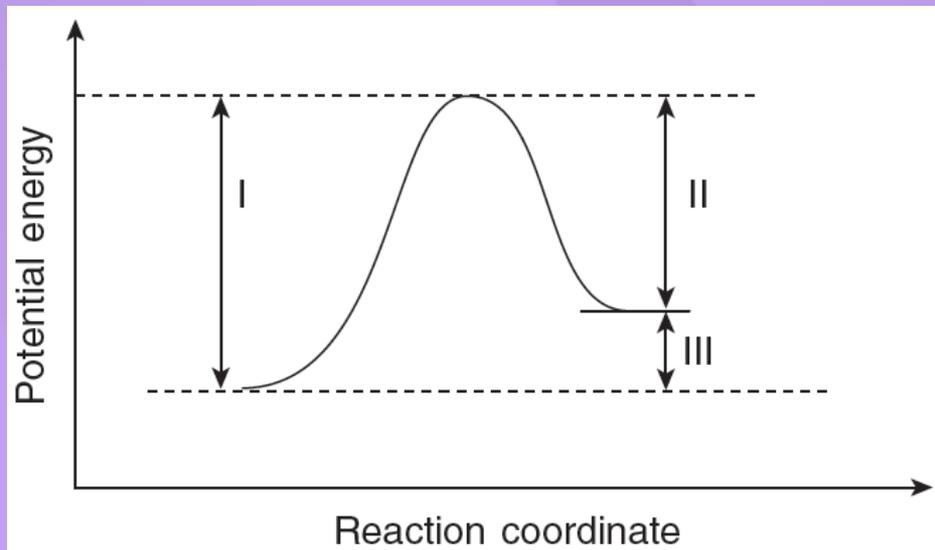
## Unit Exercise (p. 76)

- 6 In the decomposition of hydrogen peroxide, manganese(IV) oxide is a catalyst. Which of the following best describes the function of manganese(IV) oxide?
- A It speeds up the decomposition by allowing the decomposition to occur by an alternative pathway of higher activation energy.
  - B It slows down the decomposition by decreasing the collision frequency of particles.
  - C It speeds up the decomposition by increasing the average energy of the particles.
  - D It speeds up the decomposition by allowing the reaction to occur by an alternative pathway of lower activation energy.

**Answer: D**

 Unit Exercise (p. 76)

7 Which of the quantities in the potential energy profile below is / are affected by the use of a catalyst?



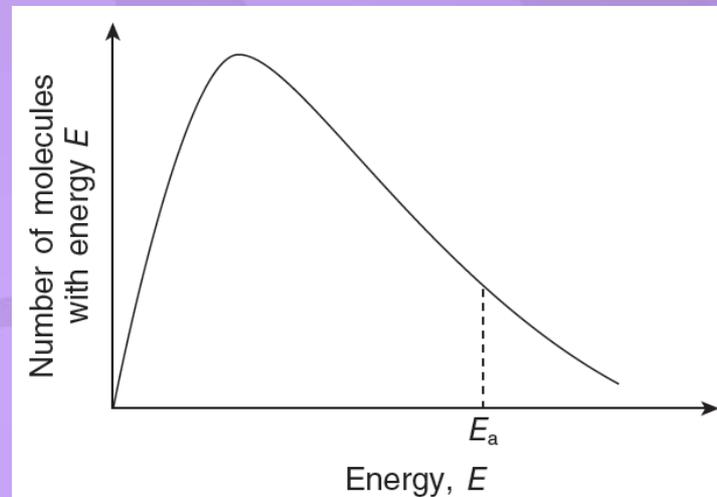
- A I only
- B III only
- C I and II only
- D II and III only

**Answer: C**

 Unit Exercise (p. 76)

Directions: Questions 8 and 9 refer to the following:

- 8  The diagram below shows the Maxwell-distribution of molecular energies for a catalysed reaction.



If the temperature were lowered, what would be the effect on the shape of the curve?

- A The peak would shift to the left and be higher.
- B The peak would shift to the left and be lower.
- C The peak would shift to the right and be higher.
- D The peak would shift to the right and be lower.

**Answer: A**

## Unit Exercise (p. 76)

9 Which of the following would shift the activation energy line to the right?

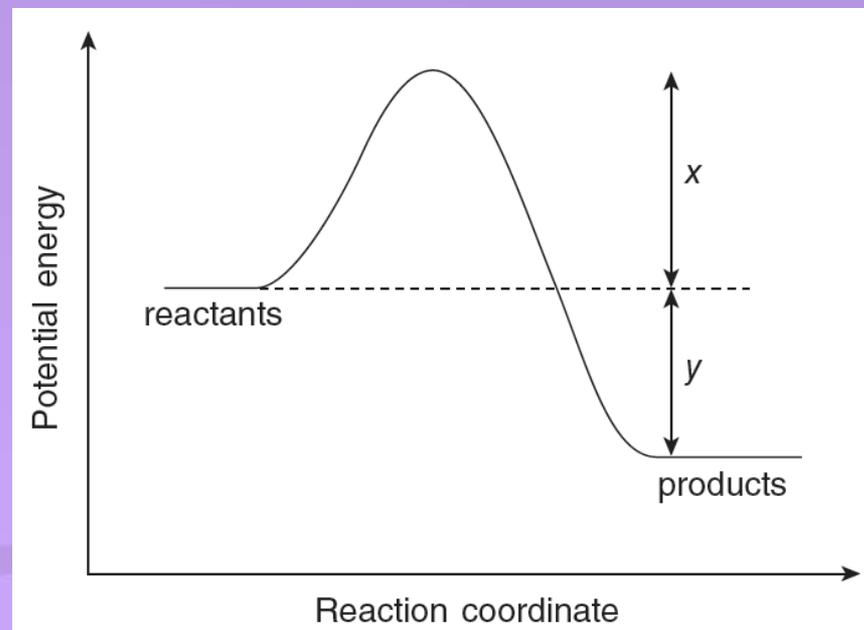
- A An increase in reactant concentration.
- B The removal of the product.
- C The removal of the catalyst.
- D The use of smaller particles with a larger surface area.

*(Edexcel Advanced Subsidiary GCE, Unit 2R, Jun. 2014, 12)*

**Answer: C**

 Unit Exercise (p. 76)

10 The potential energy profile for a reversible reaction is shown below.



Which of the following statements is / are correct?

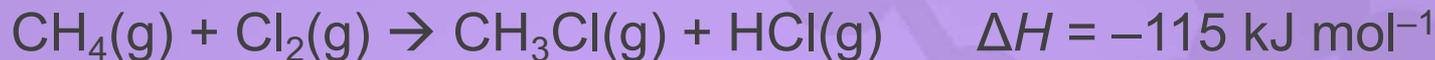
- (1) The activation energy of the forward reaction is  $x$ .
- (2) The backward reaction is endothermic.
- (3) The enthalpy change of the backward reaction is  $y - x$ .

- A (1) and (2) only  
B (1) and (3) only  
C (2) and (3) only  
D (1), (2) and (3)

**Answer: A**

 Unit Exercise (p. 76)**PART III STRUCTURED QUESTIONS**

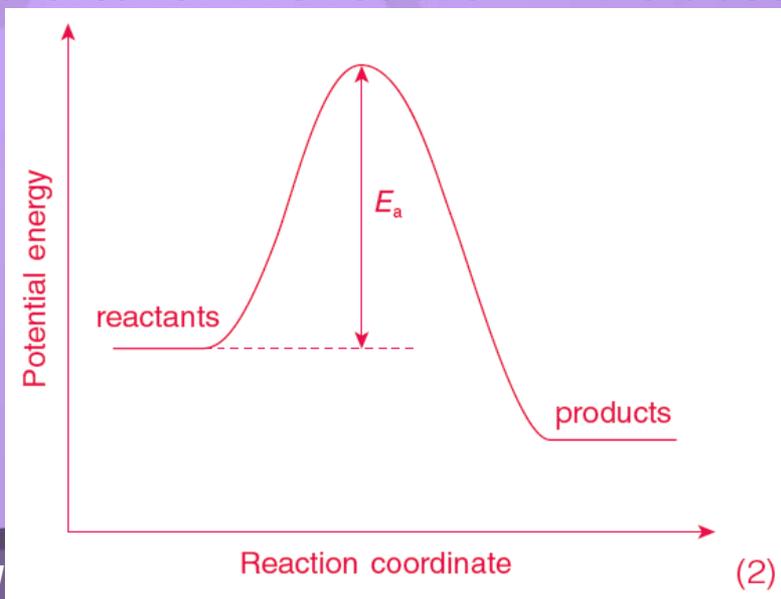
11 Methane and chlorine react according to the equation below.



a) Explain why no reaction takes place between methane and chlorine at room temperature unless the reactants are exposed to ultraviolet light or heated.

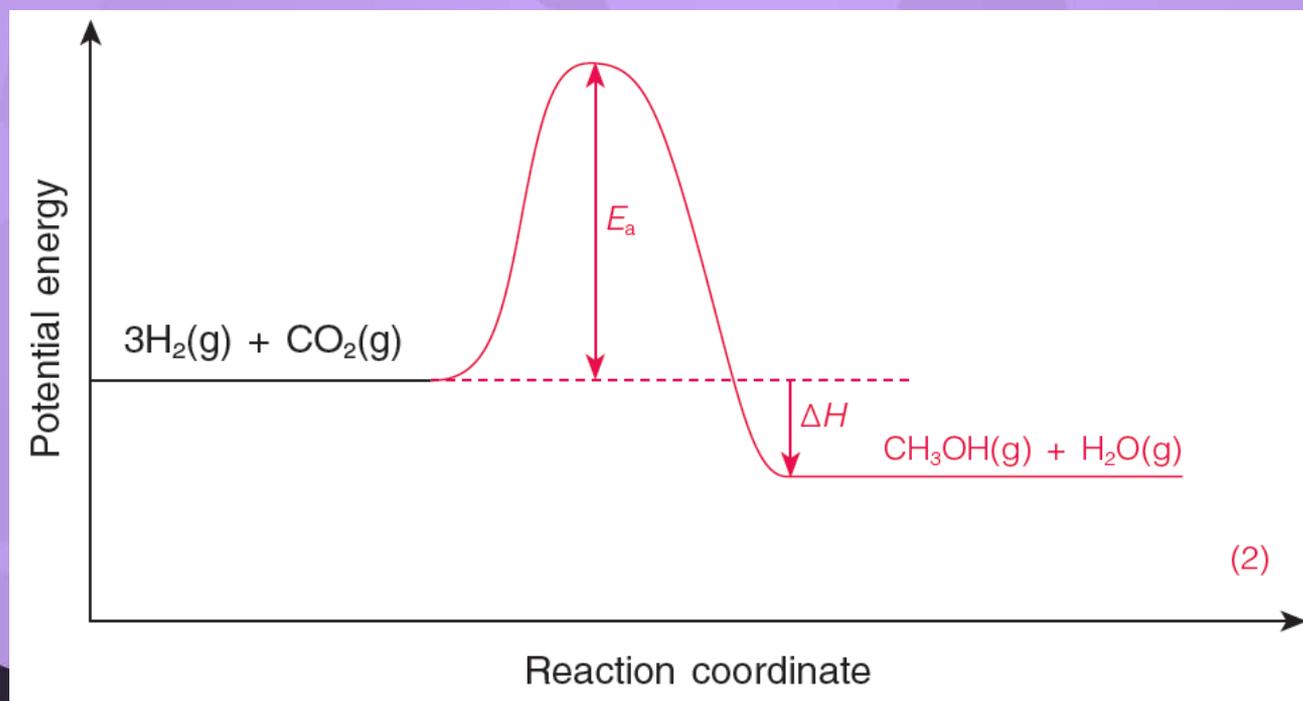
The molecules have insufficient energy to react at room temperature. (1)  
Extra energy is needed to overcome the activation energy barrier. (1)

b) Draw a potential energy profile for this reaction in the absence of light at low temperature.



 Unit Exercise (p. 76)

- 12 Methanol can be manufactured by the reaction of carbon dioxide with hydrogen:  $3\text{H}_2(\text{g}) + \text{CO}_2(\text{g}) \rightleftharpoons \text{CH}_3\text{OH}(\text{g}) + \text{H}_2\text{O}(\text{g})$   $\Delta H = -49 \text{ kJ mol}^{-1}$   
The activation energy (forward reaction) is estimated to be  $225 \text{ kJ mol}^{-1}$ .  
a) Complete the potential energy profile for the reversible reaction in the absence of any catalyst. Include the products,  $\Delta H$  and  $E_a$ .



## Unit Exercise (p. 76)

b) What is the enthalpy change of reaction for the backward reaction?

+49 kJ mol<sup>-1</sup> (1)

c) Estimate the activation energy for the backward reaction.

Activation energy for backward reaction

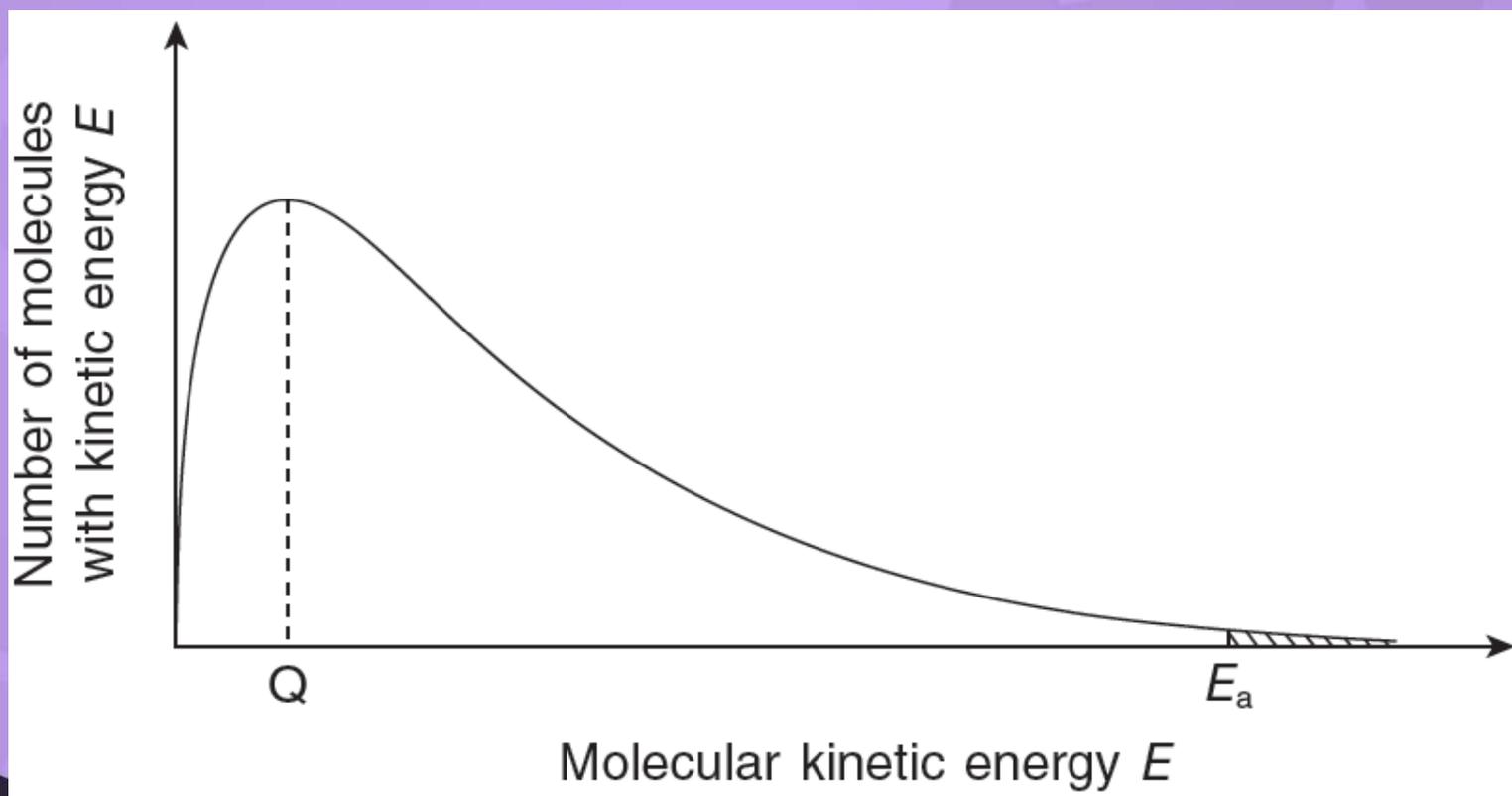
= (225 + 49) kJ mol<sup>-1</sup>

= 274 kJ mol<sup>-1</sup> (1)



## Unit Exercise (p. 76)

- 13 The graph below shows the Maxwell-Boltzmann distribution curve of molecular kinetic energies of a mixture of gases which react together. The activation energy for the reaction is  $E_a$ .





## Unit Exercise (p. 76)

- a) Explain what is meant by the term 'activation energy'  
Activation energy is the minimum kinetic energy that colliding particles must have for a collision to result in a chemical reaction. (1)
- b) State what Q represents.  
The most probable kinetic energy value (1)
- c) What does the total area under the curve represent?  
The total number of molecules in the gas (1)
- d) Why does the curve start at the origin?  
Each molecule in the system has kinetic energy greater than zero. (1)
- e) Why is the shaded area very small?  
Only a small fraction of the molecules have kinetic energy equal to or greater than the activation energy. (1)

 Unit Exercise (p. 76)

14 a) Identify TWO features of colliding molecules that react together in the gas phase.



The molecules must have sufficient energy / kinetic energy equal to or greater than the activation energy. (1)

The molecules must collide in a correct orientation. (1)

b) For many reactions, the rate approximately doubles for a 10 °C rise in temperature. Suggest why.

A temperature rise

- increases the average kinetic energy of the molecules, causing the collision frequency to increase, and hence effective collisions among molecules occur more frequently. (1)
- increases the fraction of molecules with kinetic energy equal to or greater than the activation energy, and hence a higher percentage of molecules with sufficient energy to collide and react. (1)

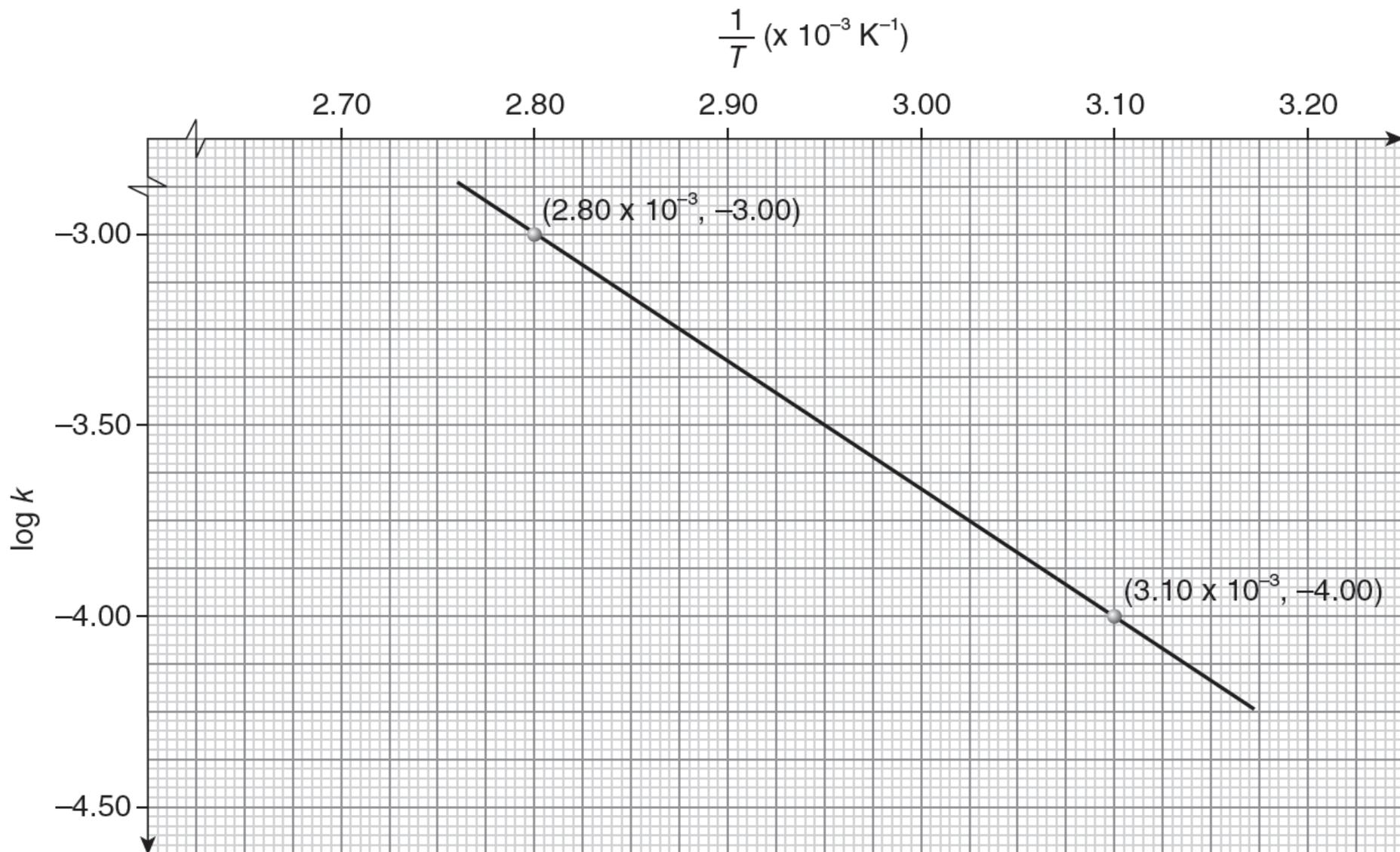


## Unit Exercise (p. 76)



15 In a chemical kinetics experiment, the rate constants ( $k$ ) of a reaction at various temperatures ( $T$ ) were determined.

The graph below shows the plot of  $\log k$  against  $\frac{1}{T}$ . Calculate the activation energy for this reaction. (Gas constant  $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ )



 Unit Exercise (p. 76)

$$\text{Slope of graph} = -\frac{E_a}{2.3R}$$

$$= \frac{-4.00 - (-3.00)}{(3.10 - 2.80) \times 10^{-3} \text{ K}}$$

$$= -3\,330 \text{ K}^{-1} (1)$$

$$E_a = 3\,330 \text{ K}^{-1} \times 2.3 \times 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$= 63\,600 \text{ J mol}^{-1}$$

$$= 63.6 \text{ kJ mol}^{-1} (1)$$

$\therefore$  the activation energy for the reaction is  $63.6 \text{ kJ mol}^{-1}$ .



## Unit Exercise (p. 76)

16 The activation energy for a certain biological process is  $50.0 \text{ kJ mol}^{-1}$ .



By what factors will the rate constant of this process increase when body temperature increases from  $37 \text{ }^\circ\text{C}$  (normal) to  $41 \text{ }^\circ\text{C}$  (fever)?

(Gas constant =  $8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ )

$$\frac{k_{41}}{k_{37}} = \frac{e^{-\frac{50\,000}{(314)(8.31)}}}{e^{-\frac{50\,000}{(310)(8.31)}}} \quad (1)$$

$$= e^{-\frac{50\,000}{(314)(8.31)} + \frac{50\,000}{(310)(8.31)}}$$

$$= e^{0.247}$$

$$= 1.28 \quad (1)$$

$\therefore$  the rate of reaction increases by a factor of 1.28.

 Unit Exercise (p. 76)

17 The rate constant of a certain reaction doubles when the temperature is  increased from 298 K to 308 K. Calculate the activation energy of the reaction.

(Gas constant  $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ ;

Arrhenius equation:  $\log k = \text{constant} - \frac{E_a}{2.3RT}$  )  
(HKDSE, Paper 2, 2015, 1(a)(iii))

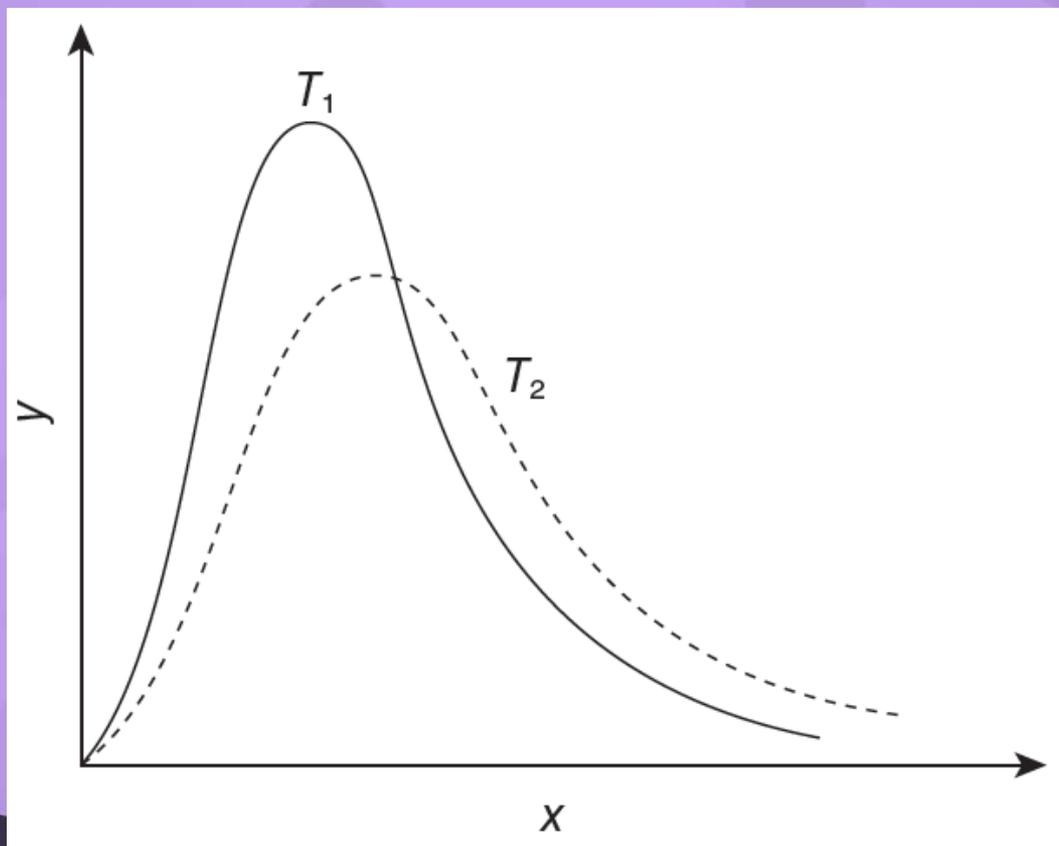
Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

 Unit Exercise (p. 76)

18 For gaseous reactions, an increase in temperature leads to an increase in reaction rate.



The graph below shows the Maxwell-Boltzmann distribution curves of molecular kinetic energies of a gas at two temperatures  $T_1$  and  $T_2$ .





## Unit Exercise (p. 76)

- What do the axes, x and y, in the above graph respectively represent?
- With reference to the above graph, suggest why an increase in temperature can lead to an increase in the reaction rate of a gaseous reaction.

*(HKDSE, Paper 2, 2013, 1(a)(i))*

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).



## Unit Exercise (p. 76)



19 Given that the activation energy for a certain biological process is  $48.0 \text{ kJ mol}^{-1}$  and the rate constant is  $2.50 \times 10^{-2} \text{ s}^{-1}$  at  $25 \text{ }^\circ\text{C}$ . What is the rate constant at  $37 \text{ }^\circ\text{C}$ ? (Gas constant =  $8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ )

$$\begin{aligned}\log \frac{k_{310}}{k_{298}} &= \frac{E_a}{2.3R} \left( \frac{1}{298} - \frac{1}{310} \right) \\ &= \frac{48\,000}{(2.3)(8.31)} \left( \frac{1}{298} - \frac{1}{310} \right) \quad (1) \\ &= 0.326 \\ \frac{k_{310}}{k_{298}} &= 2.12 \\ k_{310} &= 2.12 \times 2.50 \times 10^{-2} \text{ s}^{-1} \\ &= 0.0530 \text{ s}^{-1} \quad (1)\end{aligned}$$

$\therefore$  the rate constant at  $37 \text{ }^\circ\text{C}$  is  $0.0530 \text{ s}^{-1}$ .



## Unit Exercise (p. 76)

20 NO(g) reacts with Br<sub>2</sub>(g) according to the equation below.



The rate equation for this reaction is:

$$\text{rate} = k[\text{NO}(\text{g})]^2[\text{Br}_2(\text{g})]$$

a) A chemist lets 0.0150 mol dm<sup>-3</sup> NO(g) and 0.0200 mol dm<sup>-3</sup> Br<sub>2</sub>(g) react together.

The initial rate of this reaction is 4.78 x 10<sup>-4</sup> mol dm<sup>-3</sup> s<sup>-1</sup>.

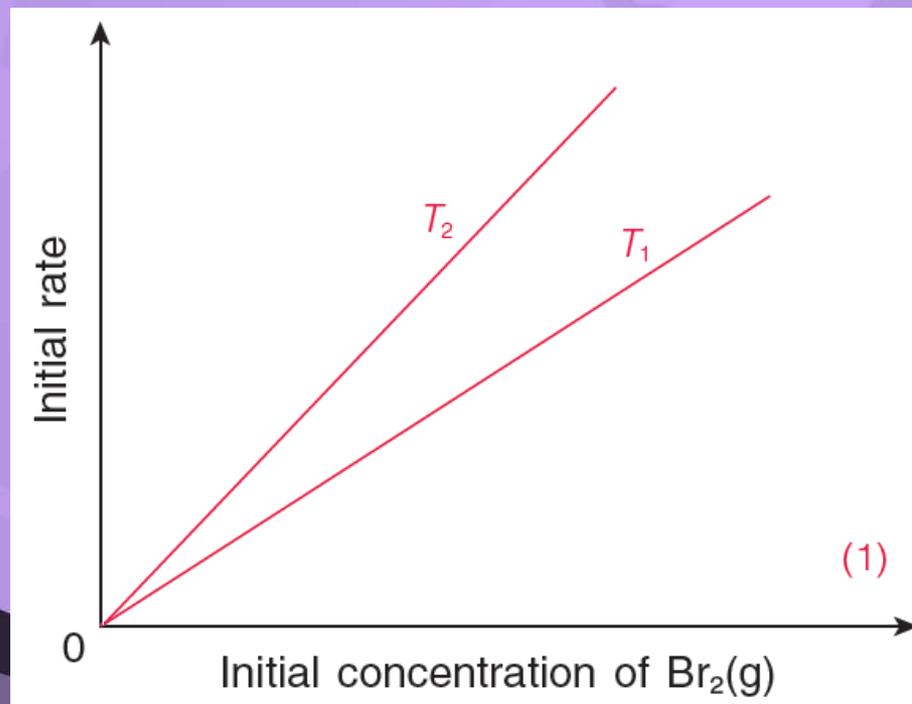
Calculate the rate constant, *k*, for this reaction. State the units, if any.

$$4.78 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1} = k(0.0150 \text{ mol dm}^{-3})^2(0.0200 \text{ mol dm}^{-3})$$

$$k = 106 \text{ dm}^6 \text{ mol}^{-2} \text{ s}^{-1} \quad (1)$$

 Unit Exercise (p. 76)

- b) Another chemist investigated the effect of changing the initial concentration of  $\text{Br}_2(\text{g})$  on the initial rate at two different temperatures,  $T_1$  and  $T_2$  using same initial concentration of excess  $\text{NO}(\text{g})$ .
- Sketch graphs of the experimental results obtained at temperature  $T_1$  and  $T_2$  where  $T_1 < T_2$ .
  - State the effect of the higher temperature on the rate constant,  $k$ .



increase

 Unit Exercise (p. 76)

21 The equation for the hydrolysis of  $\text{CH}_3\text{Br}(\text{aq})$  with  $\text{NaOH}(\text{aq})$  is shown:



The table below lists the values of the rate constants ( $k$ ) for the reaction at various temperatures ( $T$ ).

$T$ (K)	$k$ ( $\text{dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ )
293	$1.05 \times 10^{-4}$
308	$6.15 \times 10^{-4}$
323	$3.07 \times 10^{-3}$
333	$8.23 \times 10^{-3}$

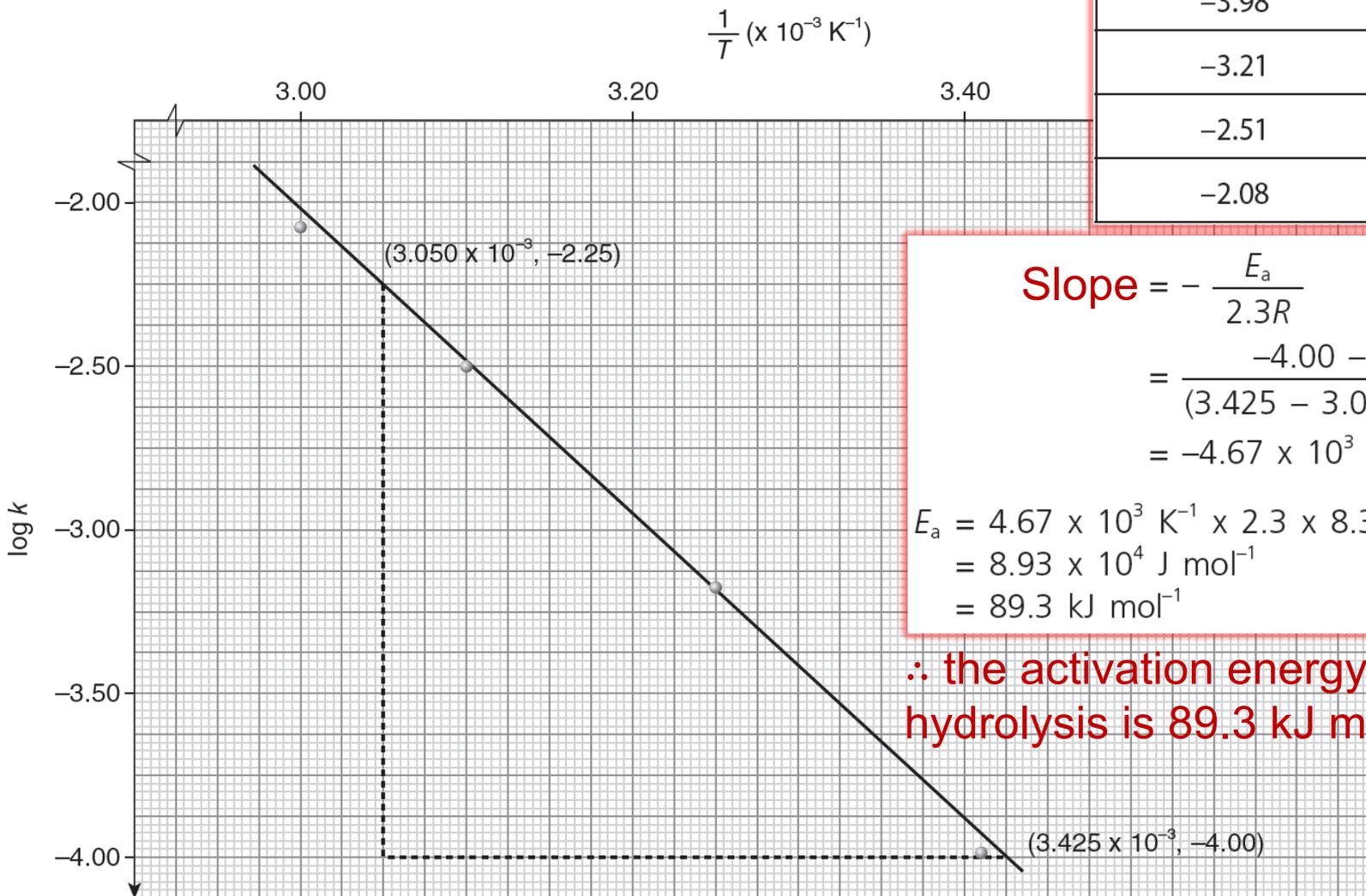
Determine the activation energy,  $E_a$ , for the hydrolysis by plotting an appropriate graph.

(Arrhenius equation:  $k = Ae^{\frac{-E_a}{RT}}$  where  $A$  is a constant and  $R$  is the gas constant ( $8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ )).



# Unit Exercise (p. 76)

$\log k$	$\frac{1}{T} (\text{K}^{-1})$
-3.98	$3.41 \times 10^{-3}$
-3.21	$3.25 \times 10^{-3}$
-2.51	$3.10 \times 10^{-3}$
-2.08	$3.00 \times 10^{-3}$



$$\begin{aligned} \text{Slope} &= -\frac{E_a}{2.3R} \\ &= \frac{-4.00 - (-2.25)}{(3.425 - 3.050) \times 10^{-3} \text{ K}} \\ &= -4.67 \times 10^3 \text{ K}^{-1} \end{aligned} \quad (1)$$

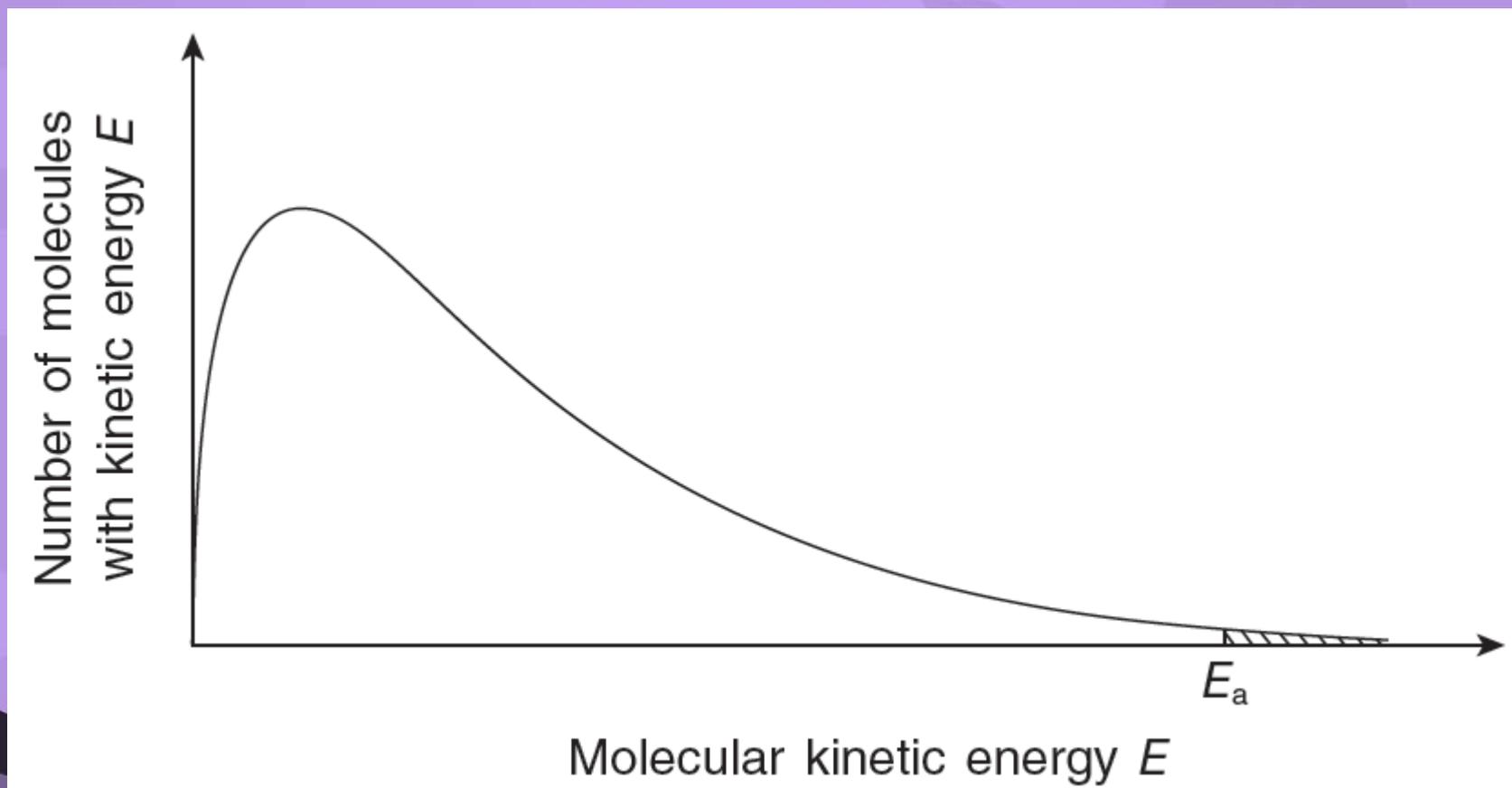
$$\begin{aligned} E_a &= 4.67 \times 10^3 \text{ K}^{-1} \times 2.3 \times 8.31 \text{ J K}^{-1} \text{ mol}^{-1} \\ &= 8.93 \times 10^4 \text{ J mol}^{-1} \\ &= 89.3 \text{ kJ mol}^{-1} \end{aligned} \quad (1)$$

$\therefore$  the activation energy for the hydrolysis is  $89.3 \text{ kJ mol}^{-1}$ .

(2)

 Unit Exercise (p. 76)

22 The graph below shows the Maxwell-Boltzmann distribution curve of  molecular kinetic energies of a gas at temperature  $T$ . The activation energy is shown by the symbol  $E_a$ .





## Unit Exercise (p. 76)

a) What does the area under the curve represent?



**The total number of molecules in the gas (1)**

b) With reference to the above graph, explain how a catalyst increases the rate of a chemical reaction.

**A catalyst increases the rate of a chemical reaction by providing an alternative reaction pathway of lower activation energy. (1)**

**With a catalyst, a greater fraction of molecules has kinetic energy equal to or greater than the lower activation energy. (1)**

c) Explain why the use of a catalyst can reduce the demand for energy.

**Allow reactions to take place at lower temperatures. (1)**

 Unit Exercise (p. 76)

23 Hydrogen peroxide decomposes to give water and oxygen.



The reaction is exothermic.

a) Explain, in terms of bond breaking and bond making, why the decomposition of hydrogen peroxide is exothermic.

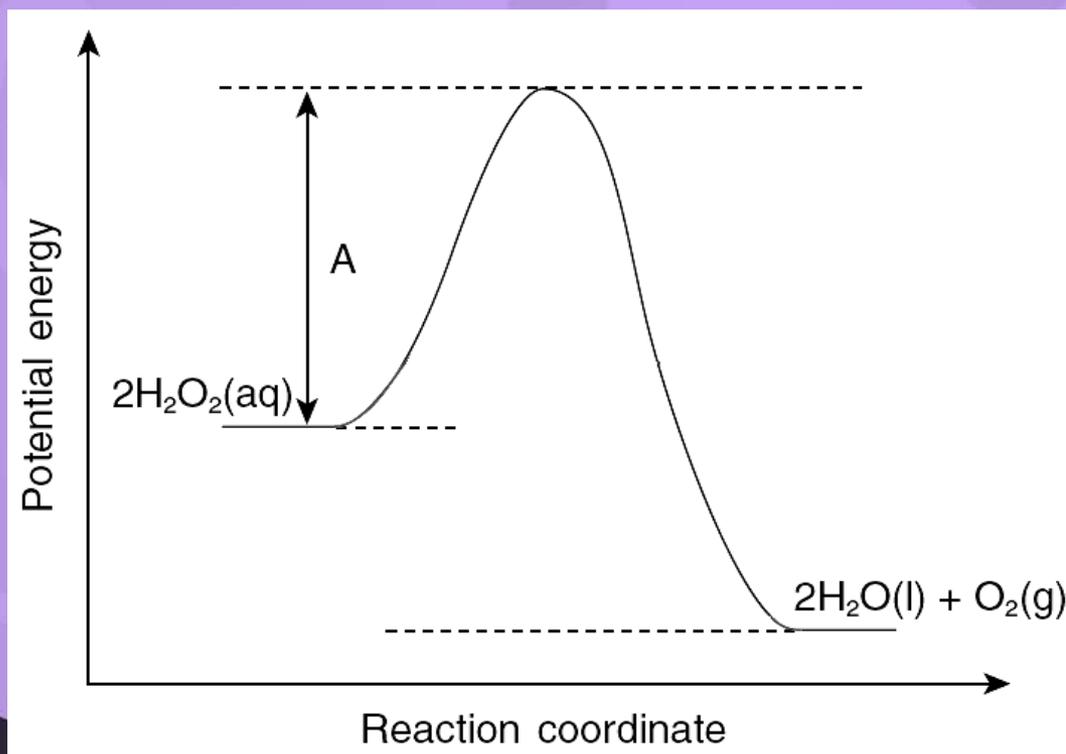
The energy required to break the existing bonds is less than the energy released from forming new bonds. (1)



## Unit Exercise (p. 76)

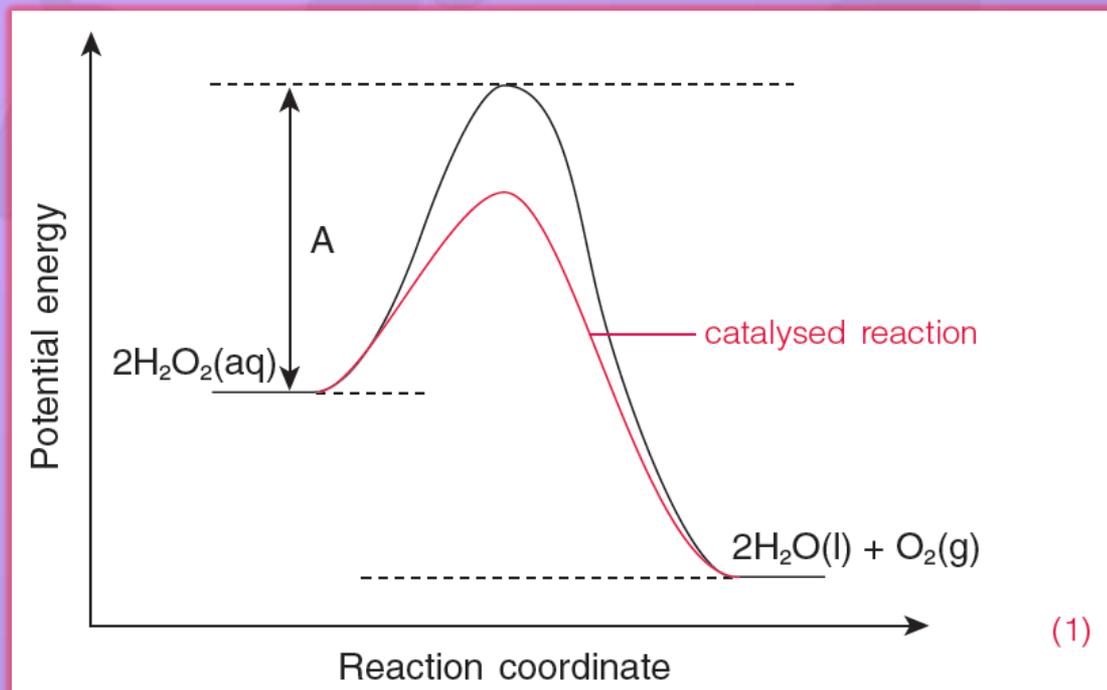
- b) The potential energy profile for this reaction is shown below. The decomposition of hydrogen peroxide is slow. What does this suggest about energy change A?

A is high. (1)



 Unit Exercise (p. 76)

- c) The decomposition of hydrogen peroxide is catalysed by manganese(IV) oxide.
- i) Explain how a catalyst increases the rate of the decomposition.  
**A catalyst increases the rate of the decomposition by providing an alternative reaction pathway of lower activation energy. (1)**
- ii) On the diagram in (b), draw the potential energy profile for the catalysed reaction.





## Unit Exercise (p. 76)

24 The decomposition of ammonia is represented by the following equation.



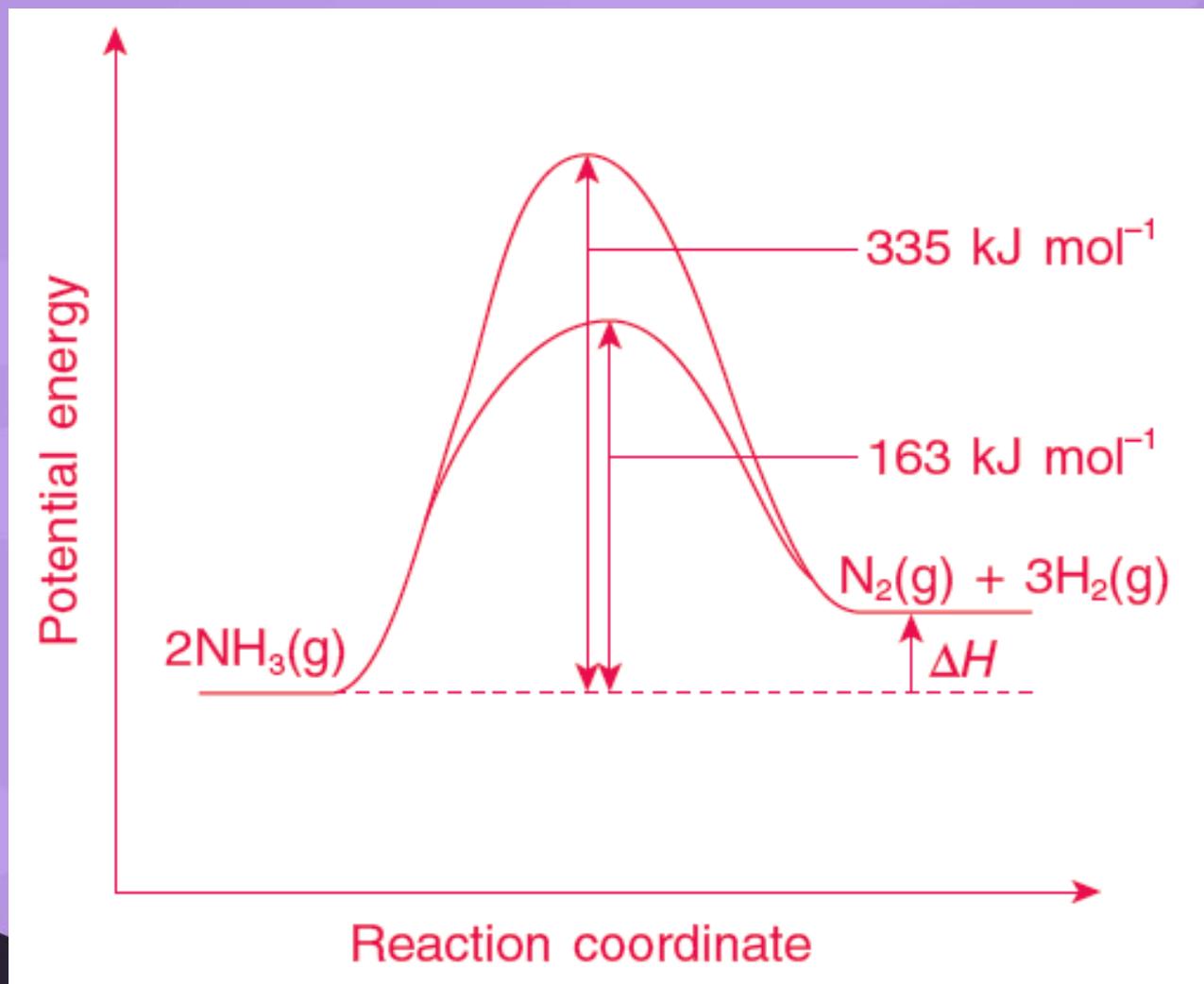
The activation energy for the uncatalysed reaction is  $335 \text{ kJ mol}^{-1}$ .  
The activation energy for the reaction when tungsten is used as a catalyst is  $163 \text{ kJ mol}^{-1}$ .

- Draw labelled potential energy profiles for the uncatalysed and catalysed reactions.
- Suppose the decomposition is carried out at  $700^\circ\text{C}$ . By what factor is the rate constant increased in the presence of a tungsten catalyst?  
(Gas constant =  $8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ )



## Unit Exercise (p. 76)

a)



(3)



## Unit Exercise (p. 76)

b)

$$\frac{k_{\text{catalysed}}}{k_{\text{uncatalysed}}} = \frac{e^{-\frac{163\,000}{(8.31)(973)}}}{e^{-\frac{335\,000}{(8.31)(973)}}} \quad (1)$$
$$= 1.7 \times 10^9 \quad (1)$$

## Unit Exercise (p. 76)

25 At room temperature, ethanol can be produced by fermentation of sugars in the presence of yeast. State the function of yeast, and explain why this function fails at high temperatures.

*(HKDSE, Paper 2, 2014, 1(a)(ii))*

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).



## Unit Exercise (p. 76)



26 Some 'green' buses use biodiesel fuel which is a fuel that has been made from waste fats and cooking oil.

The conversion of fats and oils into biodiesel needs a catalyst. The usual catalyst is hot concentrated sodium hydroxide solution.

Scientists are investigating a new catalyst. The new catalyst is an enzyme. Here is some information about both catalysts.

Feature of enzyme	Feature of hot concentrated sodium hydroxide
Speeds up reaction a lot	speeds up reaction
Easily damaged	not easily damaged
Needs warm conditions	needs hot conditions
Can be coated onto a solid surface	mixed in with the products at the end
Speeds up this reaction only	speeds up other reactions of the esters as well as this reaction
Expensive	very cheap

 Unit Exercise (p. 76)

Evaluate both catalysts. Suggest which catalyst would be best and explain why.

*(OCR GCSE (Higher Tier), Chem. A (21st Century Science), A173/02, Jun. 2015, 1(c))*

- The enzyme speeds up reaction a lot. (1)
- The enzyme is more easily damaged (i.e. enzyme becomes denatured easily). Hence the enzyme does not last as long / needs to be replaced more often. (1)
- The enzyme needs warm conditions, i.e. lower energy / costs. (1)
- The enzyme can be coated onto a solid surface. The products can be separated easily. (1)
- The enzyme speeds up this reaction only. Thus, there are fewer side reactions and hence less waste and less purification are needed. (1)
- The enzyme is expensive, but this can be justified by greater productivity. (1)

(Reverse arguments for hot concentrated sodium hydroxide are acceptable.)

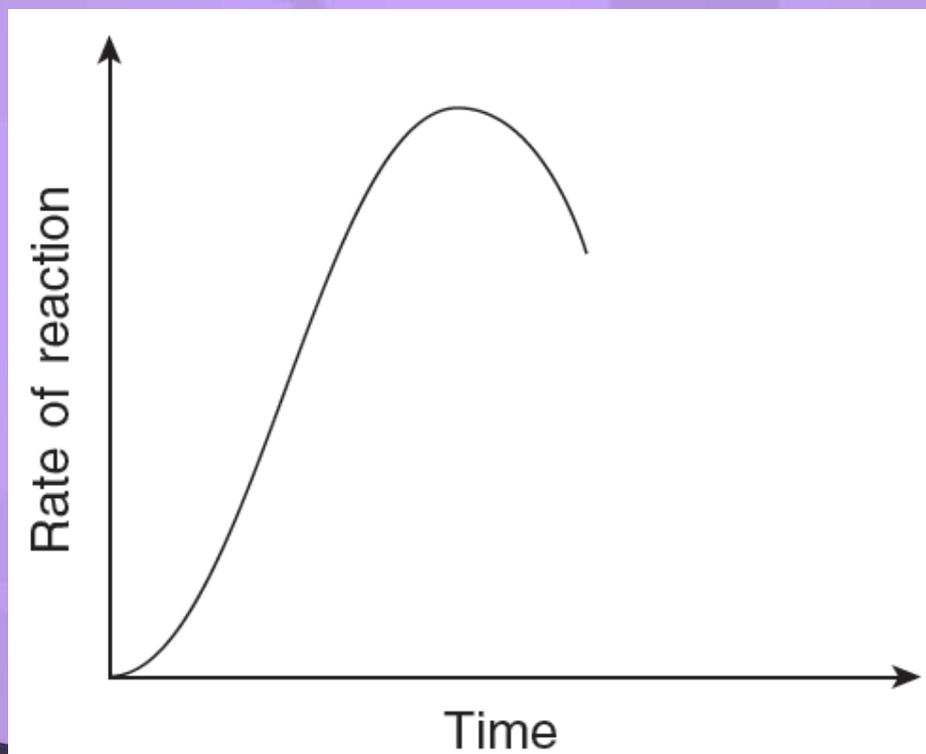
The enzyme would be best.

 Unit Exercise (p. 76)

27 The alcohol ethanol can be made by fermentation. Yeast is added to

 aqueous glucose:  $C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$

Carbon dioxide is given off and the mixture becomes warm, as the reaction is exothermic. The graph shows how the rate of reaction varies over several days.





## Unit Exercise (p.76)

a) Suggest a method of measuring the rate of this reaction.

Measure the volume of gas at regular time intervals. (1)

b) Why does the rate initially increase?

Increase in temperature / more yeast present. (1)

c) Suggest TWO reasons why the rate eventually decreases.

*(Cambridge IGCSE, 0620/04, Paper 4, Specimen paper, 2016, 8(d))*

Glucose is used up. (1)

The concentration of ethanol becomes high enough to kill the yeast. (1)