Answers to exam-style questions

Topic 6

1 C

2 A

3 B

4 C

5 C

6 C

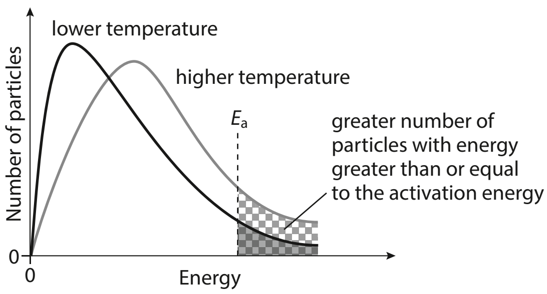
7 A

8 C

9 D

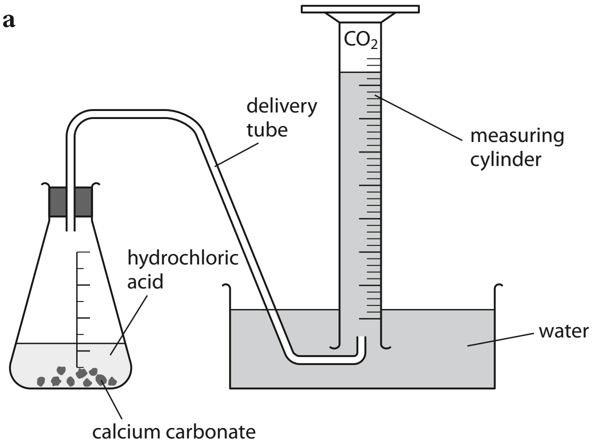
10 B

11 There are two factors that contribute to the increase in rate of reaction with increase in temperature. The first is that as the temperature increases the particles have more kinetic energy and so are moving faster; the particles collide more often. This is only, however, a small effect. The second factor is that at the higher temperature there are more particles with energy greater than the activation energy; therefore there is a greater chance that a collision will result in reaction and there will be more successful collisions per unit time. This is the more important factor in explaining why the rate of reaction increases with temperature. This can be shown on the Maxwell–Boltzman distribution – the checked area represents the number of particles with energy greater than the activation energy at the higher temperature.



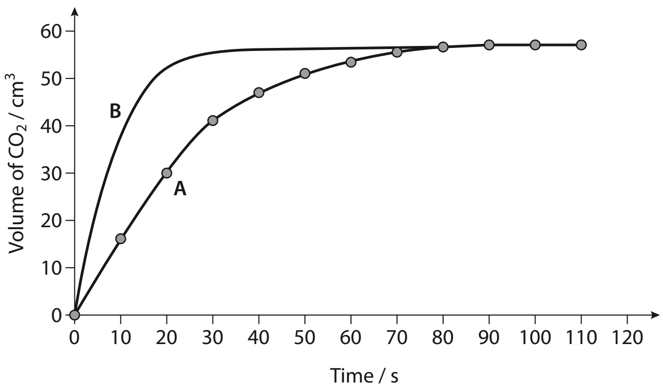
[4]

12 a



The volume of carbon dioxide produced is measured every 10 s. **[3]**

b and **d** **b[2] d[2]**



c The rate is fastest at the beginning, because the graph has its maximum gradient (slope) there. The concentration of HCl is highest at the beginning, so there are the most collisions per unit time. **[2]**

e i No. moles of $CaCO_3 = \dfrac{1.00}{100.09} = 9.99 \times 10^{-3}$ mol

No. moles of HCl $= \dfrac{20.0}{1000} \times 0.300$

$= 6.00 \times 10^{-3}$ mol

From the chemical equation: 6.00×10^{-3} mol HCl react with 3.00×10^{-3} mol $CaCO_3$. Therefore $CaCO_3$ is in excess. The number of moles of CO_2 formed is obtained by using the number of moles of the limiting reactant (HCl): 6.00×10^{-3} mol HCl produces 3.00×10^{-3} mol CO_2.

volume of $CO_2 = 3.00 \times 10^{-3} \times 24.0$

$= 0.0720\ dm^3$

There are $1000\ cm^3$ in $1\ dm^3$, so the volume in cm^3 is $72.0\ cm^3$. **[3]**

ii The most likely reason is that some gas will escape before the bung is put in the flask. **[1]**

13 a The power of a reactant's concentration in the experimentally determined rate equation. The overall order of reaction is the sum of the powers of the concentration terms in the experimentally determined rate equation. **[2]**

b From experiment 2 to 1, when the concentration of X is doubled, the rate of reaction also doubles, so the order of reaction with respect to X is 1.
From experiment 3 to 2, when the concentration of Y is doubled, the rate of reaction also doubles, so the order of reaction with respect to Y is 1.
The rate expression is rate = k[X][Y] **[4]**

c From experiment 1, [X] = 0.500 mol dm^{-3} and [Y] = 0.500 mol dm^{-3}. Substituting these values and the value of the rate into the rate expression we get:

$3.20 \times 10^{-3} = k \times 0.500 \times 0.500$

Rearranging this we get $k = 0.0128$.

The units of k are obtained by substituting units into the rate expression:

$\cancel{\text{mol dm}^{-3}}\,\text{s}^{-1} = k \times \text{mol dm}^{-3} \times \cancel{\text{mol dm}^{-3}}$

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

Rearranging this we get:

$\dfrac{\text{s}^{-1}}{\text{mol dm}^{-3}} = k$

$k = \text{mol}^{-1}\,\text{dm}^{3}\,\text{s}^{-1}$

Therefore, $k = 0.0128\,\text{mol}^{-1}\,\text{dm}^{3}\,\text{s}^{-1}$. **[2]**

d This can be worked out in two ways. In the first, the values can be substituted into the rate expression:

rate = k[X][Y]

rate = $0.0128 \times 0.100 \times 0.100$

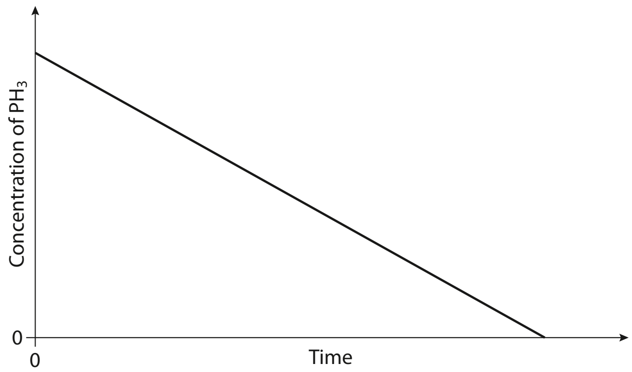
= $1.28 \times 10^{-4}\,\text{mol dm}^{-3}\,\text{s}^{-1}$

In the second method, the concentration of X is $\frac{1}{5}$ its value in experiment 1 and the reaction is first order with respect to X, so dividing the concentration of X by 5 will reduce the rate by a factor of 5; the concentration of Y is $\frac{1}{5}$ its value in experiment 1 and the reaction is first order with respect to Y, so dividing the concentration of Y by 5 will reduce the rate by a factor of 5. If we take these two factors together, the rate in the new experiment will be $\frac{1}{25}$ times the rate in experiment 1. **[2]**

rate = $3.20 \times 10^{-3} \times \frac{1}{25} = 1.28 \times 10^{-4}\,\text{mol dm}^{-3}\,\text{s}^{-1}$

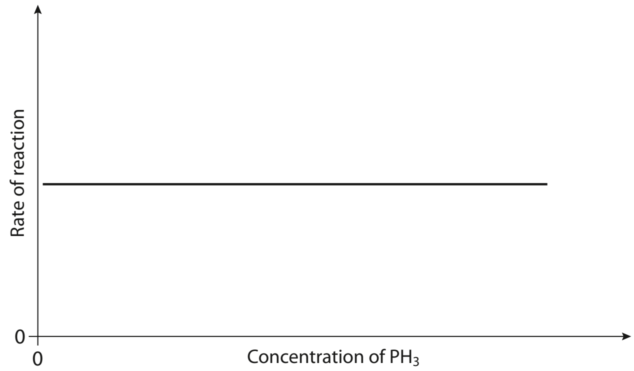
e The rate constant increases as the temperature increases. This is because more particles have energy greater than or equal to the activation energy at a higher temperature. **[2]**

14 a



[2]

b



[2]

15 a The rate equation cannot be derived directly from the stoichiometric equation because the reaction could occur in more than one step. This reaction is unlikely to occur in one step, because that would involve three molecules all colliding at the same time, which is statistically extremely unlikely. **[2]**

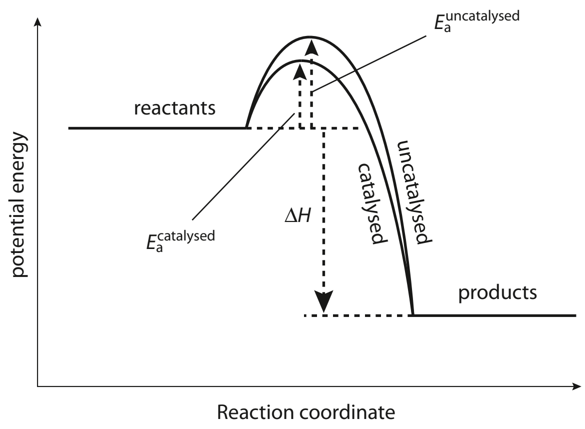
b The rate-determining step is the slowest step in a reaction mechanism: that is, the step with highest activation energy. **[1]**

c $NO + NO \rightarrow N_2O_2$ **rate-determining step**
$N_2O_2 + O_2 \rightarrow 2NO_2$ **fast**

The rate expression contains $[NO]^2$, which indicates that two molecules of NO are involved up to and including the rate-determining step – molecularity of the rate-determining step is 2 because there are two reactant molecules involved. There is no O_2 in the rate equations, which indicates that O_2 can be involved in the mechanism only in a fast step after the rate-determining step. **[4]**

d All the concentrations will be decreased by a factor of 2. The rate of the reaction depends on the concentration of NO squared, so if the concentration is halved the rate will decrease by a factor of 2^2. So the rate will be decreased by a factor of 4. **[2]**

e



$E_a^{\text{uncatalysed}}$
reactants
potential energy
$E_a^{\text{catalysed}}$
ΔH
catalysed
uncatalysed
products
Reaction coordinate

[4]