# Kinetics

## Rate determining step

**1.** For each of the everyday processes described below, identify the step that slows the process down.

1. Making a cup of tea

A close-up of a sign

Description automatically generated

1. Playing with a model helicopter received as a Christmas present

A black and white image of a person

Description automatically generated

1. Getting out of the house in the morning on time

A white background with black and white clouds

Description automatically generated

(3 marks)

The overall rate of these processes is controlled by the *rate of the slowest step*. For a chemical reaction we call this step the **rate determining** or **rate limiting** **step**.

A close-up of a diagram

Description automatically generatedFor each of the multi-step reactions below, write the overall equation for the reaction and identify the rate limiting step.

**BONUS MARK**

In a chemical reaction, any step that occurs after the rate determining step will not affect the rate. Therefore any species that are involved in the mechanism after the rate determining step do not appear in the rate expression. Use this information to predict which of the options below is the correct rate expression for the reaction shown in question **2**.

(a) Rate = *k* [CH3Br] or (b) Rate = *k* [CH3Br][OH–]

(1 mark)

## Calculating reaction rate

1. What is the definition for *the rate of a reaction*?

(2 marks)

1. A simple way to determine the rate of a reaction is to measure the change in concentration of one reagent with time. The graph below shows the change in concentration of [C4H9Cl] during the reaction;

C4H9Cl + OH– → C4H9OH + Cl–

Graph of a graph with a blue line

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A graph paper with a graph

Description automatically generated

# Measuring reaction rate in the lab

The rate of a reaction is defined as *the change in concentration of reactants or products per unit time.* The units of rate are mol dm−3 s−1.

The method chosen to measure the rate of a reaction depends on the individual reaction.

For each of the reactions below, use the observations made to **calculate the initial rate of the reaction**.

1. Measuring the rate of a reaction when a precipitate is formed;

Na2S2O3(aq) + 2 HCl(aq) → 2 NaCl(aq) + H2O(l) + S(s) + SO2(g)

A student wished to investigate how temperature affected the rate of the reaction between sodium thiosulfate and acid. He reacted 10 cm3 of a 0.02 mol dm−3 solution of sodium thiosulfate with 40 cm3 of hydrochloric acid (excess) at 22 °C. The time taken to produce a precipitate of 1 × 10–4 mol of sulfur was found to be 56 s.

Initial rate of production of sulfur = mol dm–3 s–1

(2 marks)

1. Measuring the rate of a reaction in which there is a change in colour;

CH3COCH3(aq) + I2(aq) → CH3COCH2I(aq) + H+(aq) + I− (aq)

colourless brown colourless colourless

A student followed the reaction between iodine and propanone to produce iodopropanone. She set up the first experiment as described in the table below and found it took 279 s for the brown colour of the iodine to disappear.

A close-up of a graph

Description automatically generated

1. Measuring the rate of a reaction in which a gas is produced;

Mg(s) + 2 HCl(aq) → MgCl2(aq) + H2(g)

The student reacted a 3 cm strip of magnesium ribbon with 25 cm3 of 2.0 mol dm–3 HCl (an excess). He found that 14 cm3 of gas was produced in the first 10 seconds of the reaction.

(You may assume the reaction was carried out at RTP where 1 mole of gas has a volume of 24 dm3.)

Initial rate of loss of hydrochloric acid = mol dm–3 s–1

(4 marks)

# Determining the rate equation

For each of the following sets of experimental data determine;

(a) The rate equation for the reaction,

A table of equations with numbers and numbers

Description automatically generated with medium confidence(b) The value of the rate constant, *k* including its units.

# Arrhenius and rate

**1.** A reaction between A and B was found to be first order with respect to both.

(a) Write the rate equation for the reaction

(1 mark)

(b) The rate constant *k* varies with temperature. Use the data together with the rate equation you have written in part (a) to complete the table below;

(4 marks)

A table with numbers and symbols

Description automatically generated

A graph paper with a graph and a graph on it

Description automatically generated(c) A scientist wishes to use this data to determine the activation energy for the reaction. He can do this using the Arrhenius equation;

# Kinetics – Answers

## Rate determining step

**1.** (a) Step 3 Boil the water

(b) Step 3 Charge the batteries for 24 h

(c) Step 2 Get out of bed (although this may depend on the individual!)

*(3 marks)*

**2.** *Overall equation:* CH3Br + OH– → CH3OH + Br–

*Rate limiting step:* CH3Br → CH3+ + Br− (Step 1)

*(2 marks)*

**3.** *Overall equation:* 2 NO + O2 → 2 NO2

*Rate limiting step:* NO + NO → N2O2 (Step 1)

*(2 marks)*

**4.** *Overall equation:* 2 NO + 2 H2 → N2 + 2 H2O

*Rate limiting step:* N2O2 + H2 → N2O + H2O (Step 2)

*(2 marks)*

**BONUS MARK** Answer = (a) Rate = *k* [CH3Br]

*(1 mark)*

## Calculating reaction rate

1. The rate of a reaction is the change in concentration of reactants or products per unit time *(2 marks)*

2. (a) i. 2.0 × 10–4 mol dm–3 s–1

ii. 1.3 × 10–4 mol dm–3 s–1

iii. 5.5 × 10–5 mol dm–3 s–1

*(1 mark for each correct value, 1 mark for the correct units for rate)*

(b)

A graph with a line graph

Description automatically generated

1. *marks)*

(c) The reaction is first order with respect to C4H9Cl.

*(1 mark)*

# Measuring reaction rate in the lab

**1.** Change in concentration of sulfur = (1 × 10–4 mol – 0 mol) / 0.05 dm3 = 2 × 10–3 mol dm–3

*(1 mark)*

Initial rate of production of sulfur = 2 × 10–3 mol dm–3 /56 s = 3.6 × 10–5 mol dm–3 s–1

*(1 mark)*

**2.** Moles of iodine in reaction mixture at start = 0.002 dm3 × 0.005 mol dm–3 = 1 × 10–5 mol

*(1 mark)*

Change in concentration of iodine = (1 × 10–5 mol – 0 mol) / 0.025 dm3 = 4 × 10–4 mol dm–3

*(1 mark)*

Initial rate of loss of iodine = 4 × 10–4 mol dm–3 /279 s = 1.4 × 10–6 mol dm–3 s–1

*(1 mark)*

**3.** Moles in 14 cm3 of hydrogen at RTP = 0.014 dm3 / 24 dm3 = 5.8 × 10–4 mol

*(1 mark)*

Moles of acid used up to produce this many moles of hydrogen = 5.8 × 10–4 mol × 2 = 1.17 × 10–3 mol

*(1 mark)*

Change in concentration of acid = 1.17 × 10–3 mol / 0.025 dm3 = 0.047 mol dm–3

*(1 mark)*

Initial rate of loss of hydrochloric acid = 0.047 mol dm–3 / 10 s = 0.0047 mol dm–3 s–1

*(1 mark)*

## Determining the rate equation

**1.** Rate = *k*[A]2; *k* = 280 mol–1 dm3 s–1

*(2 marks for the identification of the correct order wrt A and B, 1 mark for k with correct units)*

**2.** Rate = *k*[B]; *k* = 2.5 × 10–3 s–1

*(2 marks for the identification of the correct order wrt A and B, 1 mark for k with correct units)*

**3.** Rate = *k*[X]2[Y]; *k* = 2.08 mol–2 dm6 s–1

*(3 marks for the identification of the correct order wrt X, Y and Z, 1 mark for k with correct units)*

# Arrhenius and rate

**1.** (a) Rate = *k*[A][B]

*(1 mark)*

A table with numbers and symbols

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(c)

A graph with numbers and a line

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