6.2.2 Clock reactions and activation energy

Name: _____ Class: _____ Date: _____ **Total: 19 marks**

KEY WORDS clock reaction • activation energy • initial rate

Objective

Run the clock-reaction experiment and its arithmetic: the thiosulfate's role, why volumes are held identical, completing the reciprocal-and-log table, and the activation energy from the Arrhenius gradient. Original practice for WCH16 Q3 (Summer 2025; CP9b/CP10).

1 Key knowledge and worked examples

■ What the thiosulfate does

It reacts with iodine as it forms ($2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$), **delaying the colour change** until a fixed small amount of iodine has been produced; then the starch flashes **colourless → blue-black**. Same waiting amount every run, so $1/t \propto$ initial rate. Not: "to stop the reaction" – the scheme ignores quenching language.

■ Why a burette for every solution

Keeping the **total volume identical** keeps the reactant concentrations identical between runs – the credited reason. "More accurate" is a fallback, not the answer. Both tubes equilibrate to the water bath before mixing; record the temperature of the mixture.

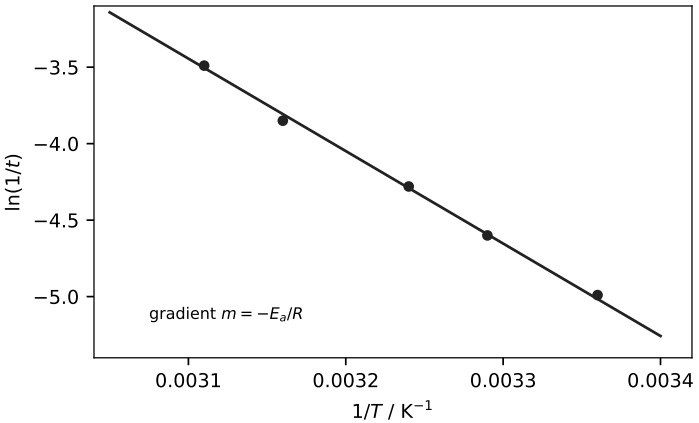
■ Table discipline

$T(\text{K}) = T(^\circ\text{C}) + 273$; $1/T$ to 3 s.f.; $1/t$ to 3 s.f.; $\ln(1/t)$ to 2 d.p. Check pairings row by row – a 316 K row belongs to 43 °C, not 49 °C.

■ The gradient pays in kJ mol^{-1}

$$\ln(1/t) = -\frac{E_a}{R} \cdot \frac{1}{T} + \text{constant}$$

Gradient $m = -E_a/R$, so $E_a = -mR = 6045 \times 8.31 = 50\,234\ \text{J mol}^{-1} = 50.2\ \text{kJ mol}^{-1}$ – positive, kJ mol^{-1}, sign carried by the gradient.



2 Practice

2.1 Explain the role of thiosulfate ions in the clock reaction, in two points the scheme credits. [2]

2.2 Why is the starch delivered from a burette rather than a measuring cylinder, even though its exact concentration does not matter? [2]

2.3 Complete: at 31 °C ($T = 304$ K, $t = 99$ s), calculate $1/t$ and $\ln(1/t)$. [2]

2.4 A graph of $\ln(1/t)$ against $1/T$ has gradient -7210 K. Calculate E_a in kJ mol^{-1}. [2]

3 Exam-style questions

3.1 For the peroxodisulfate–iodide clock: state the colour change at the endpoint; explain why both boiling tubes stand in the water bath until their temperatures match before mixing; and state what is plotted to give a straight line. [4]

3.2 The table gives 49 °C → 322 K → 0.00311 K^{-1} → t = 33 s, and 25 °C → 298 K → 0.00336 K^{-1} → t = 147 s. Complete both rows' $1/t$ and $\ln(1/t)$ values, and state the factor by which the rate changes over this range. [4]

3.3 A student's plotted gradient is -6045 K with $R = 8.31$ J mol^{-1} K^{-1}. Calculate E_a, and explain why quoting the answer as -50.2 kJ mol^{-1} loses the mark. [3]

Solutions

Original-task marking: M1 method, A1 accuracy, B1 reasoning.

2.1 The thiosulfate reacts with/removes the iodine as it forms, delaying the colour change until a fixed small amount of iodine has been produced – so $1/t$ measures the initial rate.

2.2 A burette delivers the same volume each run, keeping the total volume – and so all the reactant concentrations – constant between temperatures.

2.3 $1/t = 1/99 = 0.0101\ \text{s}^{-1}$; $\ln(0.0101) = -4.60$.

2.4 $E_a = 7210 \times 8.31 = 59\,916\ \text{J mol}^{-1}$; $= 59.9\ \text{kJ mol}^{-1}$.

3.1 Colourless $\rightarrow$ blue-black. Equilibrating means the reaction starts at the intended temperature – mixing cold solutions would change the temperature being tested. Plot $\ln(1/t)$ against $1/T$.

3.2 49 °C: $1/t = 1/33 = 0.0303$, $\ln = -3.49$ (or -3.50); 25 °C: $1/t = 1/147 = 0.00680$, $\ln = -4.99$. Rate comparison: $0.0303/0.00680 = 4.5$ – the rate at 49 °C is about 4.5 times the rate at 25 °C.

3.3 $E_a = -(-6045) \times 8.31 = 50\,234\ \text{J mol}^{-1} = 50.2\ \text{kJ mol}^{-1}$. The negative sign belongs to the gradient; E_a is defined positive – a negative activation energy is physically meaningless.