Arrhenius and activation energy

OCR A H432 · Full A-level

Show your working where needed. Check the worked answers when you finish.

Name _____ Date _____

1. Explain the effect of temperature

The same reaction mixture is tested at two temperatures. Both runs have the same initial concentrations. Predict how the rate constant changes at the higher temperature. Explain your prediction.

2. Calculate a rate constant

Use $k = A\exp(-E_a/RT)$. $A = 4.20 \times 10^{8}\ s^{-1}$, $E_a = 61.0\ kJ\ mol^{-1}$, $T = 320\ K$ and $R = 8.31\ J\ K^{-1}\ mol^{-1}$. Calculate k.

3. Rearrange for activation energy

Use $k = A\exp(-E_a/RT)$. For a first-order reaction, $k = 0.0450\ s^{-1}$, $A = 3.00 \times 10^{10}\ s^{-1}$ and $T = 335\ K$. $R = 8.31\ J\ K^{-1}\ mol^{-1}$.

Calculate E_a in $kJ\ mol^{-1}$.

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4. An Arrhenius plot

A plot of ln k against 1/T has a gradient of -7.30×10^3 K.

Use $\ln k = -E_a/(RT) + \ln A$ and R = 8.31 J K^{-1} mol^{-1} to calculate E_a.

5. Two points on a best-fit line

Two points on an ln k against 1/T line are (0.00290 K^{-1}, −2.80) and (0.00340 K^{-1}, −6.55). Use $\ln k = -E_a/(RT) + \ln A$ and R = 8.31 J K^{-1} mol^{-1}.

Calculate E_a. Use your gradient and either point to determine A. The units of k are s^{-1}.

6. Predict k at a second temperature

A first-order reaction has $k_1 = 0.00640$ s^{-1} at $T_1 = 300$ K and $E_a = 50.0$ kJ mol^{-1}.

Use $\ln(k_2/k_1) = (E_a/R)(1/T_1 - 1/T_2)$, with R = 8.31 J K^{-1} mol^{-1}, to find k_2 at 320 K.

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7. Compare two pathways

Two ln k against 1/T plots have gradients of −9000 K and −6200 K. Use $\ln k = -E_a/(RT) + \ln A$.

Identify the pathway with the lower E_a. Is the gradient alone sufficient to decide which pathway has the larger k at 300 K? Explain.

8. Rate constants at two temperatures

A first-order reaction has $k_1 = 0.0260\ s^{-1}$ at 310 K and $k_2 = 0.0920\ s^{-1}$ at 332 K.

Use $\ln(k_2/k_1) = (E_a/R)(1/T_1 - 1/T_2)$, with $R = 8.31\ J\ K^{-1}\ mol^{-1}$, to calculate E_a in kJ mol^{-1}.

9. Extra challenge · Check a measured rate constant

Two points on an Arrhenius line are ($1/T = 0.00320\ K^{-1}$, $\ln k = -5.00$) and ($1/T = 0.00350\ K^{-1}$, $\ln k = -7.40$). Here k is in s^{-1}.

Calculate E_a and A, then find the factor by which a measured k of $0.0900\ s^{-1}$ at 340 K exceeds the predicted k. Use $\ln k = -E_a/(RT) + \ln A$ and $R = 8.31\ J\ K^{-1}\ mol^{-1}$.

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Worked answers

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1. Explain the effect of temperature

The rate constant increases. A larger fraction of particles has energy at least equal to the activation energy, so more collisions are successful per second. At fixed concentrations, the rate equation attributes the higher rate to a larger k.

2. Calculate a rate constant

Convert E_a to 61 000 J mol^{-1}. Then k = (4.20 × 10^{8}) exp[−61 000 ÷ (8.31 × 320)] ≈ 0.0458 s^{-1}.

3. Rearrange for activation energy

Taking natural logarithms gives $E_a = RT \ln(A/k)$.

E_a = 8.31 × 335 × ln[(3.00 × 10^{10})/0.0450] = 75 800 J mol^{-1}, or 75.8 kJ mol^{-1} to three significant figures.

4. An Arrhenius plot

The slope is $-E_a/R$. E_a = −(−7.30 × 10^{3}) × 8.31 = 60 663 J mol^{-1}.

To three significant figures, E_a = 60.7 kJ mol^{-1}.

5. Two points on a best-fit line

Gradient = (−6.55 + 2.80) ÷ (0.00340 − 0.00290) = −7500 K. Thus E_a = 7500 × 8.31 = 62.3 kJ mol^{-1}.

The intercept is −2.80 − (−7500 × 0.00290) = 18.95. Since ln A = 18.95, A = exp(18.95) ≈ 1.70 × 10^{8} s^{-1}.

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Worked answers

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6. Predict k at a second temperature

$\ln(k_2/k_1) = (50\,000 \div 8.31)(1/300 - 1/320) = 1.2535\ldots$

$k_2 = 0.00640 \times \exp(1.2535\ldots) = 0.0224\ \text{s}^{-1}$.

7. Compare two pathways

The −6200 K slope has the smaller magnitude, so it corresponds to the lower E_a.

The intercept ln A is also needed. At a fixed 1/T the value of ln k depends on both slope and intercept.

8. Rate constants at two temperatures

$E_a = R \ln(k_2/k_1) \div (1/T_1 - 1/T_2)$. Substitute 8.31 ln(0.0920/0.0260) ÷ (1/310 − 1/332).

This gives 49 100 J mol^{-1}, or 49.1 kJ mol^{-1} to three significant figures.

9. Extra challenge · Check a measured rate constant

Gradient = −2.40/0.00030 = −8000 K. Thus $E_a = 8000 \times 8.31 = 66.5$ kJ mol^{-1}. The intercept ln A = −5.00 + 8000(0.00320) = 20.6, so $A = \exp(20.6) = 8.84 \times 10^8\ \text{s}^{-1}$.

At 340 K, ln k = 20.6 − 8000/340, so predicted $k = 0.0534\ \text{s}^{-1}$. The measured value is 0.0900/0.0534 = 1.68 times the predicted value.

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