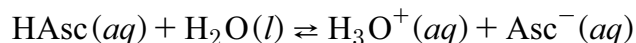
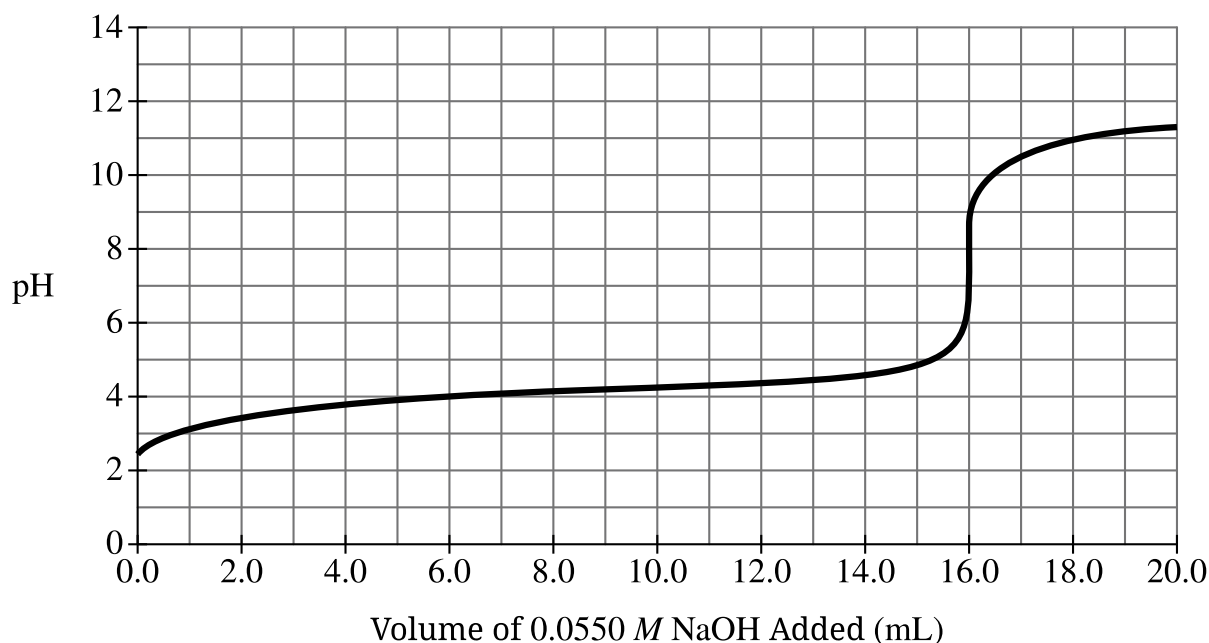


2. Answer the following questions about ascorbic acid (vitamin C).

- A.** A student combusts a sample of ascorbic acid, $C_xH_yO_z$, to determine its chemical composition. The only products of the reaction are 0.2400 mol of CO_2 and 2.883 g of H_2O .
- Calculate the number of moles of H_2O produced.
 - The mole ratio of carbon (C) to oxygen (O) is 1:1 in ascorbic acid. Based on this information and your answer to part A (i), determine the empirical formula of ascorbic acid.
- B.** Ascorbic acid, $HAsc(aq)$, acts as a weak acid, as shown in the equation.



The following titration curve was produced when a 10.0 mL sample of $HAsc(aq)$ was titrated using 0.0550 M $NaOH(aq)$.



- Calculate the molar concentration of the ascorbic acid solution.
- From the titration curve, determine the approximate pK_a of ascorbic acid.
- What is the value of the ratio $\frac{[Asc^-]}{[HAsc]}$ when the pH of the solution is 4.7?

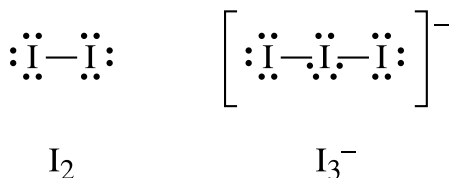
- C. Dehydroascorbic acid (DHAsc) can be produced by reacting ascorbic acid with the triiodide ion, I_3^- , as represented by the following equation.



The student runs three trials of the reaction with different initial concentrations of HAsc and I_3^- , producing the following data.

Trial	[HAsc] (M)	$[\text{I}_3^-]$ (M)	Initial Rate of DHAsc Formation (M/s)
1	0.450	1.200	2.457×10^{-4}
2	0.450	0.600	1.229×10^{-4}
3	0.900	1.200	4.914×10^{-4}

- The rate law for the reaction is $\text{rate} = k[\text{HAsc}][\text{I}_3^-]$. Explain how the data in the table support the conclusion that the reaction is first order with respect to [HAsc].
 - Calculate the value of the rate constant, k , for the reaction. Include units with your answer.
- D. The triiodide ion, I_3^- , is significantly more soluble in water than elemental iodine, I_2 , is. Identify an intermolecular force between I_3^- and water that is **not** present between I_2 and water, which could explain the difference in solubility. Lewis diagrams for I_2 and I_3^- are provided.



Answer the following questions about an experiment in which $\text{CaCO}_3(s)$ is combined with $\text{HCl}(aq)$, represented by the following balanced equation.



(a) Write the balanced net ionic equation for the reaction.

A student performs an investigation to study factors that affect the rate of the reaction. In each trial the student combines 50.0 mL of $\text{HCl}(aq)$ at 21.2°C with 1.00 g of $\text{CaCO}_3(s)$ and measures the time required for the reaction to go to completion. The data are given in the following table.

Trial	Concentration of $\text{HCl}(aq)$ (M)	Particle Size of $\text{CaCO}_3(s)$	Time of Reaction (s)
1	1.00	Fine powder	67
2	1.00	Small chunks	112
3	1.00	Large chunk	342
4	3.00	Fine powder	22
5	3.00	Small chunks	227
6	3.00	Large chunk	114

(b) The student correctly identifies that trial 5 is inconsistent with the other trials. Explain why the student's claim is correct using the data in the table.

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Continue your response to **QUESTION 3** on this page.

- (c) Based on the reaction conditions and the collisions that occur between particles, explain the reason for the difference in the reaction times for trial 2 and trial 3.
- (d) The student claims that the reaction is zero order with respect to $\text{HCl}(aq)$. Do you agree or disagree with the student's claim? Justify your answer using the student's data.
- (e) The $\text{HCl}(aq)$ was present in excess in all trials of the experiment. Determine the molarity of the $\text{HCl}(aq)$ in the beaker after the reaction is complete in trial 2. Assume that the volume of the mixture remains constant at 50.0 mL throughout the trial. (The molar mass of CaCO_3 is 100.09 g/mol.)

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Continue your response to Question 5 on this page.



In order to measure the enthalpy of the reaction shown, the student repeats trial 1 by mixing 50.0 mL of $\text{HCl}(aq)$ with 1.00 g of $\text{CaCO}_3(s)$ using a coffee cup calorimeter. The student records the temperature of the system every 20 seconds. The data are given in the following table.

Time (s)	Measured Temperature of Solution ($^{\circ}\text{C}$)
0	21.20
20	21.51
40	21.70
60	21.85
80	21.90
100	21.90

(f) Is the reaction endothermic or exothermic? Justify your answer using the information in the table.

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Continue your response to **QUESTION 3** on this page.

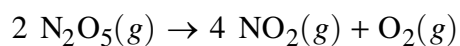
(g) Based on the experimental data, the mass of the system is 51.0 g, and the specific heat of the reaction mixture is $4.0 \text{ J} / (\text{g} \cdot ^\circ\text{C})$.

(i) Calculate the magnitude of heat transfer, q , in joules.

(ii) Calculate the enthalpy of reaction in units of kJ/mol_{rxn} . Include the algebraic sign on your answer.

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. The following equation represents the decomposition of N_2O_5 , for which the rate law is $\text{rate} = k[\text{N}_2\text{O}_5]$.

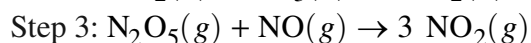
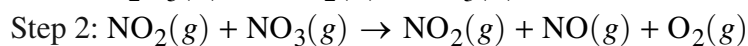
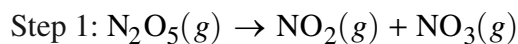


A sample of pure $\text{N}_2\text{O}_5(g)$ is placed in an evacuated container and allowed to decompose at a constant temperature of 300 K. The concentration of $\text{N}_2\text{O}_5(g)$ in the container is measured over a period of time, and the measurements are recorded in the following table.

Time (hr)	$[\text{N}_2\text{O}_5](M)$
0	0.160
1.67	0.0800
3.33	0.0400
5.00	0.0200

(a) Determine the value of the rate constant, k , for the reaction. Include units in your answer.

(b) The following mechanism is proposed for the decomposition of $\text{N}_2\text{O}_5(g)$.



Identify which step of the proposed mechanism (1, 2, or 3) is the rate-determining step. Justify your answer in terms of the rate law given.

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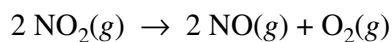
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Continue your response to **QUESTION 5** on this page.

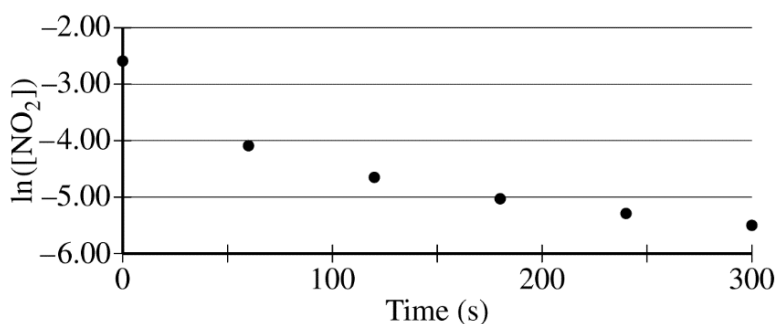
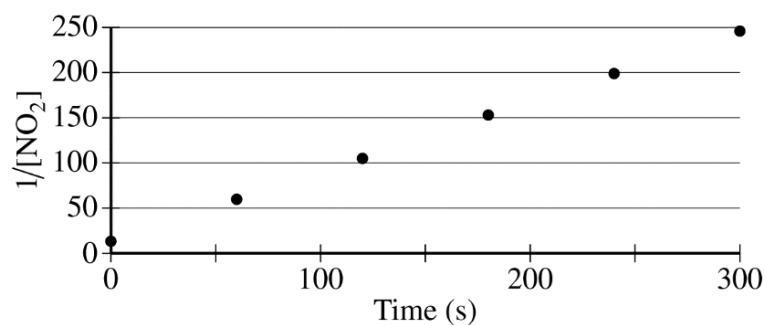
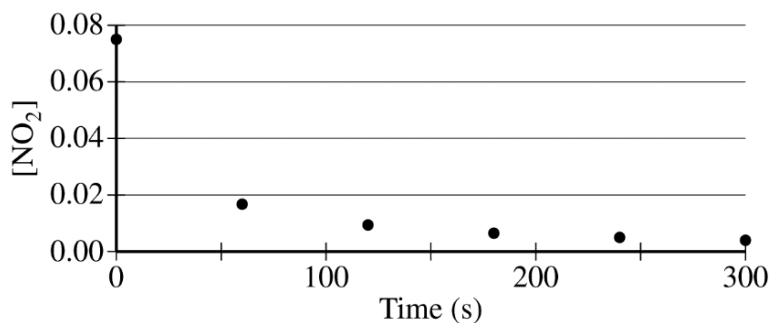
- (c) If this experiment was repeated at the same temperature but with twice the initial concentration of N_2O_5 , would the value of k increase, decrease, or remain the same? Explain your reasoning.

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6. Nitrogen dioxide, $\text{NO}_2(g)$, is produced as a by-product of the combustion of fossil fuels in internal combustion engines. At elevated temperatures $\text{NO}_2(g)$ decomposes according to the equation below.



The concentration of a sample of $\text{NO}_2(g)$ is monitored as it decomposes and is recorded on the graph directly below. The two graphs that follow it are derived from the original data.

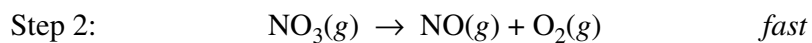
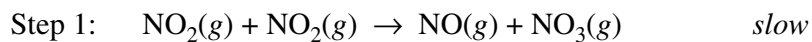


- (a) Explain how the graphs indicate that the reaction is second order.
- (b) Write the rate law for the decomposition of $\text{NO}_2(g)$.

(c) Consider two possible mechanisms for the decomposition reaction.

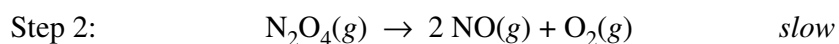
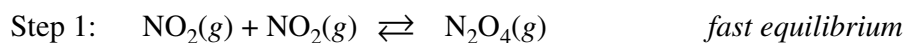
- (i) Is the rate law described by mechanism I shown below consistent with the rate law you wrote in part (b) ? Justify your answer.

Mechanism I



- (ii) Is the rate law described by mechanism II shown below consistent with the rate law you wrote in part (b) ? Justify your answer.

Mechanism II



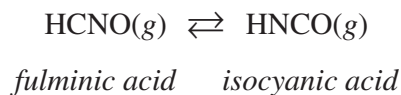
2. Answer the following questions about the isomers fulminic acid and isocyanic acid.

Two possible Lewis electron-dot diagrams for fulminic acid, HCNO, are shown below.



- (a) Explain why the diagram on the left is the better representation for the bonding in fulminic acid. Justify your choice based on formal charges.

Fulminic acid can convert to isocyanic acid according to the equation below.



Fulminic Acid	Isocyanic Acid
$\text{H}-\text{C}\equiv\text{N}-\ddot{\text{O}}:$	$\text{H}-\ddot{\text{N}}=\text{C}=\ddot{\text{O}}:$

- (b) Using the Lewis electron-dot diagrams of fulminic acid and isocyanic acid shown in the boxes above and the table of average bond enthalpies below, determine the value of ΔH° for the reaction of HCNO(g) to form HNCO(g).

Bond	Enthalpy (kJ/mol)		Bond	Enthalpy (kJ/mol)		Bond	Enthalpy (kJ/mol)
N–O	201		C=N	615		H–C	413
C=O	745		C≡N	891		H–N	391

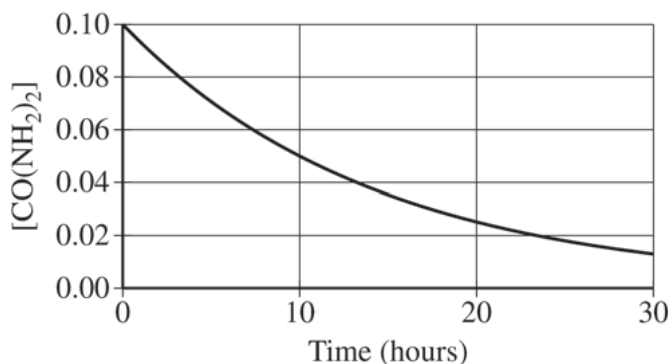
- (c) A student claims that ΔS° for the reaction is close to zero. Explain why the student's claim is accurate.
- (d) Which species, fulminic acid (HCNO) or isocyanic acid (HNCO), is present in higher concentration at equilibrium at 298 K? Justify your answer in terms of thermodynamic favorability and the equilibrium constant.

The ammonium salt of isocyanic acid is a product of the decomposition of urea, $\text{CO}(\text{NH}_2)_2$, represented below.

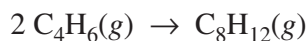


A student studying the decomposition reaction runs the reaction at 90°C . The student collects data on the concentration of urea as a function of time, as shown by the data table and the graph below.

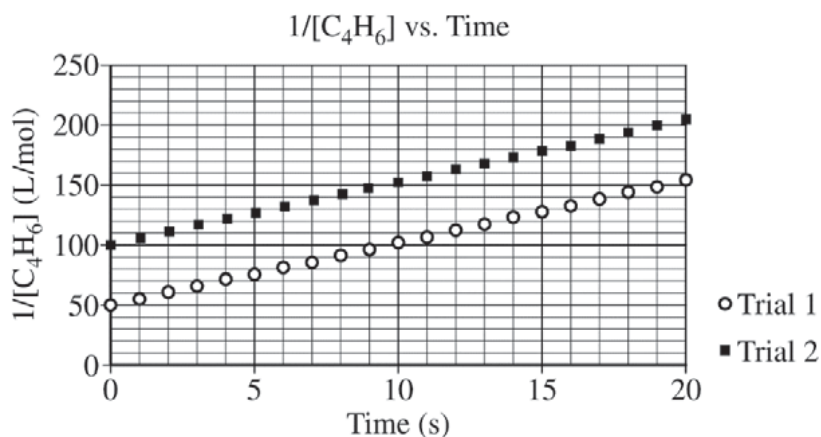
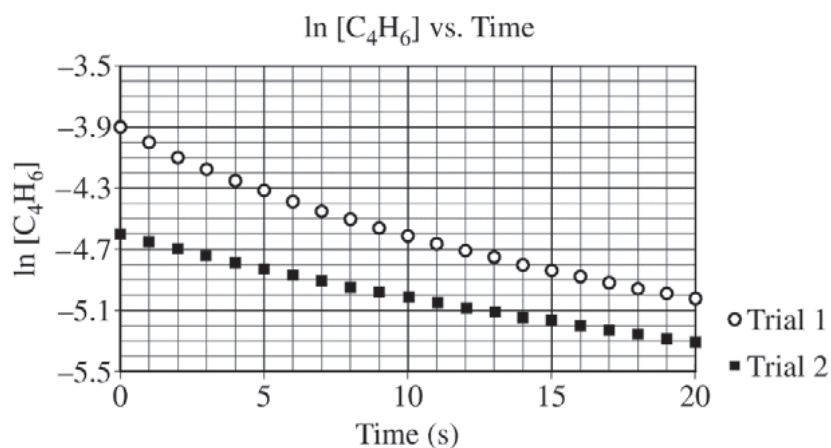
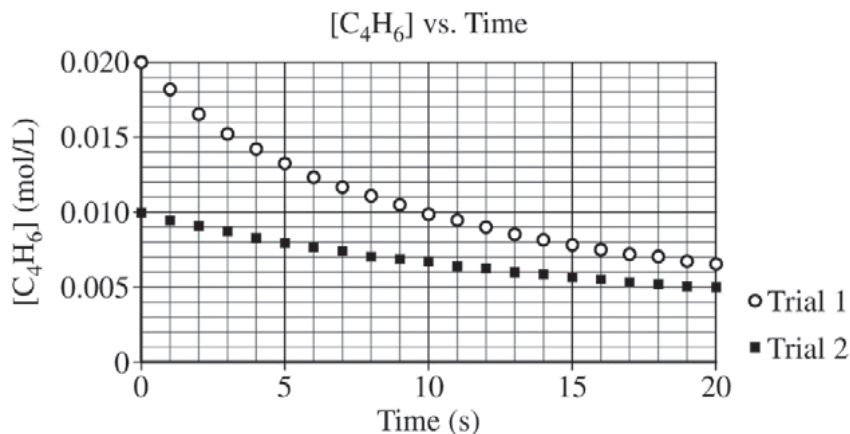
Time (hours)	$[\text{CO}(\text{NH}_2)_2]$
0	0.1000
5	0.0707
10	0.0500
15	0.0354
20	0.0250
25	0.0177
30	0.0125



- (e) The student proposes that the rate law is $\text{rate} = k[\text{CO}(\text{NH}_2)_2]$.
- Explain how the data support the student's proposed rate law.
 - Using the proposed rate law and the student's results, determine the value of the rate constant, k . Include units with your answer.
- (f) The student learns that the decomposition reaction was run in a solution with a pH of 13. Briefly describe an experiment, including the initial conditions that you would change and the data you would gather, to determine whether the rate of the reaction depends on the concentration of $\text{OH}^-(aq)$.

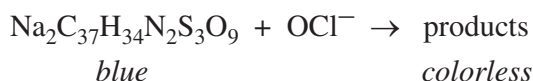


5. At high temperatures the compound C_4H_6 (1,3-butadiene) reacts according to the equation above. The rate of the reaction was studied at 625 K in a rigid reaction vessel. Two different trials, each with a different starting concentration, were carried out. The data were plotted in three different ways, as shown below.

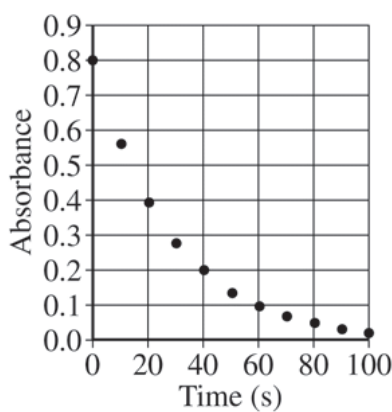


- (a) For trial 1, calculate the initial pressure, in atm, in the vessel at 625 K. Assume that initially all the gas present in the vessel is C_4H_6 .
- (b) Use the data plotted in the graphs to determine the order of the reaction with respect to C_4H_6 .
- (c) The initial rate of the reaction in trial 1 is $0.0010 \text{ mol}/(\text{L}\cdot\text{s})$. Calculate the rate constant, k , for the reaction at 625 K.

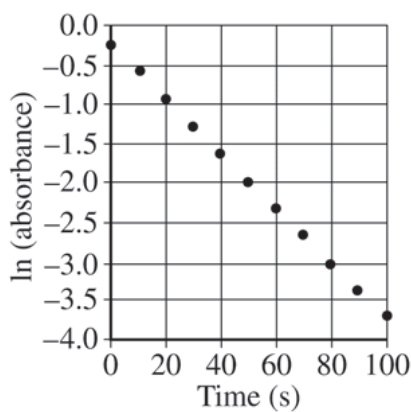
AP Chemistry: 2015



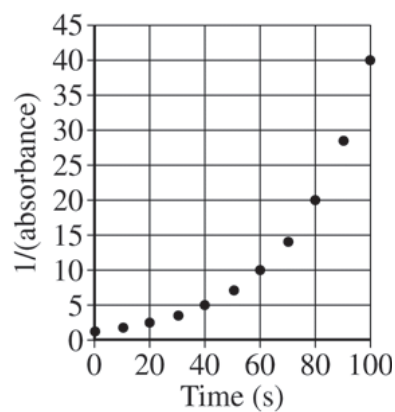
5. Blue food coloring can be oxidized by household bleach (which contains OCl^-) to form colorless products, as represented by the equation above. A student used a spectrophotometer set at a wavelength of 635 nm to study the absorbance of the food coloring over time during the bleaching process. In the study, bleach is present in large excess so that the concentration of OCl^- is essentially constant throughout the reaction. The student used data from the study to generate the graphs below.



Graph I



Graph II

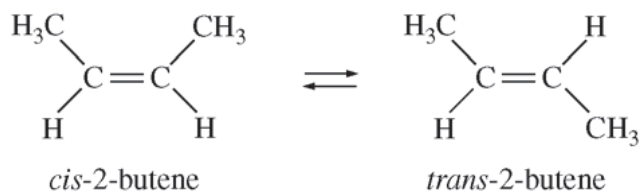


Graph III

- (a) Based on the graphs above, what is the order of the reaction with respect to the blue food coloring?
- (b) The reaction is known to be first order with respect to bleach. In a second experiment, the student prepares solutions of food coloring and bleach with concentrations that differ from those used in the first experiment. When the solutions are combined, the student observes that the reaction mixture reaches an absorbance near zero too rapidly. In order to correct the problem, the student proposes the following three possible modifications to the experiment.
- Increasing the temperature
 - Increasing the concentration of the food coloring
 - Increasing the concentration of the bleach

Circle the one proposed modification above that could correct the problem, and explain how that modification increases the time for the reaction mixture to reach an absorbance near zero.

- (c) In another experiment, a student wishes to study the oxidation of red food coloring with bleach. How would the student need to modify the original experimental procedure to determine the order of the reaction with respect to the red food coloring?



7. The half-life ($t_{1/2}$) of the catalyzed isomerization of *cis*-2-butene gas to produce *trans*-2-butene gas, represented above, was measured under various conditions, as shown in the table below.

Trial Number	Initial $P_{\text{cis-2-butene}}$ (torr)	V (L)	T (K)	$t_{1/2}$ (s)
1	300.	2.00	350.	100.
2	600.	2.00	350.	100.
3	300.	4.00	350.	100.
4	300.	2.00	365	50.

- (a) The reaction is first order. Explain how the data in the table are consistent with a first-order reaction.
- (b) Calculate the rate constant, k , for the reaction at 350. K. Include appropriate units with your answer.
- (c) Is the initial rate of the reaction in trial 1 greater than, less than, or equal to the initial rate in trial 2? Justify your answer.
- (d) The half-life of the reaction in trial 4 is less than the half-life in trial 1. Explain why, in terms of activation energy.

STOP

END OF EXAM