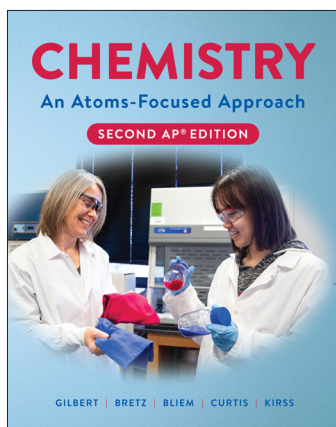




CHEMISTRY

AN ATOMS-FOCUSED APPROACH

SECOND AP® EDITION

Thomas R. Gilbert, *Northeastern University* • **Stacey Lowery Bretz**, *Ohio Northern University* • **Carribeth Bliem**, *University of North Carolina at Chapel Hill* • **Anna C. Curtis**, *University of North Carolina at Chapel Hill* • **Rein V. Kirss**, *Northeastern University* • **Todd Abronowitz**, *Parish Episcopal School* • **Paul Bonvallet**, *The College of Wooster* • **Kristen Jones**, *A&M Consolidated High School (Retired)*

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AP® CHEMISTRY TEST BANK SAMPLE

ABOUT THE TEST BANK

The AP® Test Bank contains 275 AP®-style multiple-choice questions based on data sets, equations, and molecular structures, as well as 62 AP® free-response questions. All AP®-style questions are classified by AP® Learning Objective, Science Practice, and Topic Number. The test bank also includes approximately 2,500 traditional multiple-choice and 800 short-answer questions for content recall assessment that are classified by section, difficulty, learning objective, and Bloom's Taxonomy.

ABOUT THE TEST BANK AUTHORS

Todd Abronowitz has taught chemistry since 1990 and currently teaches at Parish Episcopal School in Dallas, Texas. For all but five of those years, he has taught AP® Chemistry. He received his BS from Eastern Michigan University and his MS in Chemistry from Texas A&M University-Commerce. Since 2003 he has conducted College Board workshops, training teachers across the country. He has received numerous awards including NMSI All-American Teacher of the Year, Radio Shack National Teacher Award, ACT2 Outstanding Chemistry Teacher in Texas, Wal-Mart Teacher of the Year and the Shultz Award from the American Chemical Society. He was also featured in the book *Top Texas Teachers* as one of the top 35 educators in the state of Texas.

Kristen Jones retired from teaching AP® and Pre-AP® Chemistry after 26 years at A&M Consolidated High School in College Station, TX. She has a BS from South Dakota State University and MS from Texas A&M University, both in Food Sciences. Kristen has been teaching AP®/ Pre-AP® science teacher workshops across the US for the past 27 years for the College Board. She enjoys working with students and teachers in multiple ways to make them more successful. Kristen has received the American Chemistry Council Catalyst award, Siemens Foundation award for AP®, American Chemical Society Southwest Regional award, and College Board AP® Special Recognition award. She was named the Outstanding Secondary Science Teacher in Texas and the ACT2 Outstanding Chemistry Teacher in Texas. Her students had a >95% pass rate on the AP® exam, and they have also been winners in many national science competitions.

ABOUT NORTON TESTMAKER

Norton Testmaker brings Norton's high-quality testing materials online. Create assessments for your course from anywhere with an Internet connection, without downloading files or installing specialized software. Search and filter test bank questions by chapter, type, difficulty, AP® Learning Objectives, and other criteria. You can also edit questions or create your own. Easily export your tests to Microsoft Word or Common Cartridge files for import into your LMS.

Casey Akin is a member of the Fightin' Texas Aggie Class of 2002 and has taught Pre-AP® and AP® Chemistry for the past 17 years. She is a teacher at College Station High School in College Station, Texas. She has been honored each year as a district Star Educator and a Hall of Fame Teacher by the College Station Education Foundation and was the 2015–16 teacher of the year at her school. She has been a College Board consultant for 10 years, a National Math and Science Initiative consultant for four years, and an editor and question writer for various textbook companies for the past six years. She enjoys coaching her students in their academic competitions such as BEST Robotics, Chemistry and Physics Olympiads, UIL Academics in science, Intel ISEF and National Science Bowl.

Michael Farabaugh is a high school chemistry teacher at Albemarle High School in Charlottesville, VA. He has a BS in chemistry from The College of William and Mary, a MS in chemistry from Michigan State University, and a MT (Master of Teaching) from the University of Virginia. He has been teaching high school chemistry since 1998. He has been a reader for the AP® Chemistry Exam since 2011. He has been an item writer for the AP® Chemistry exam since 2014. He is currently a member of the AP® Chemistry Development Committee. Michael strives to create a classroom environment that includes dynamic instruction and collaborative learning. He enjoys developing assessments that address higher-order cognitive skills.



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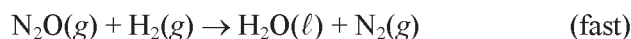
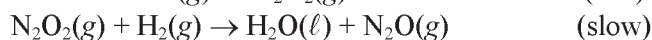
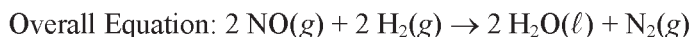
Below is a sample of the various AP® question types from Chapter 9: Properties of Gases: The Air We Breathe. If you need to see a greater sampling of assessment questions during your textbook review, please contact your Norton representative.

AP® MULTIPLE-CHOICE QUESTIONS

10 AP® multiple-choice questions are included per chapter and are based on data sets, equations, and molecular structures.

EXAMPLE:

The formation of nitrogen gas from nitrogen monoxide may occur through a three-step mechanism. This potential mechanism, as well as the overall equation, is shown.



Which of the following is the correct rate law for this reaction?

- a. $\text{Rate} = k[\text{NO}]^2$
- b. $\text{Rate} = k[\text{N}_2\text{O}_2][\text{H}_2]$
- c. $\text{Rate} = k[\text{NO}]^2[\text{H}_2]$
- d. $\text{Rate} = k \frac{[\text{N}_2\text{O}_2]^2 [\text{H}_2]}{[\text{NO}]^2}$

ANS: C

To find the rate law, we must first focus on the slow step of the mechanism, using the reactants and their coefficients to write a rate law.

$$\text{Rate} = k[\text{N}_2\text{O}_2][\text{H}_2]$$

This rate law includes a reaction intermediate, N_2O_2 , which cannot be included in the rate law. To find what to replace it with, use the following steps:

Assign rate symbols to the forward and reverse reaction for the fast step.

$$k_f[\text{NO}]^2 = k_r[\text{N}_2\text{O}_2]$$

Isolate the intermediate.

$$[\text{N}_2\text{O}_2] = [\text{NO}]^2$$

Substitute this into the original expression in place of the intermediate.

$$\text{Rate} = k[\text{NO}]^2[\text{H}_2]$$

The various rate constants in this mechanism are combined into the final k value.

DIF: Difficult

OBJ: AP Learning Objective: TRA-5.C Identify the rate law for a reaction from a mechanism in which the first step is not rate limiting.

TOP: AP Topic: 5.9

KEY: AP

MSC: AP Practice: 5.B



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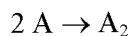
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AP® SHORT FREE-RESPONSE QUESTIONS

2 AP® short free-response questions are included per chapter.

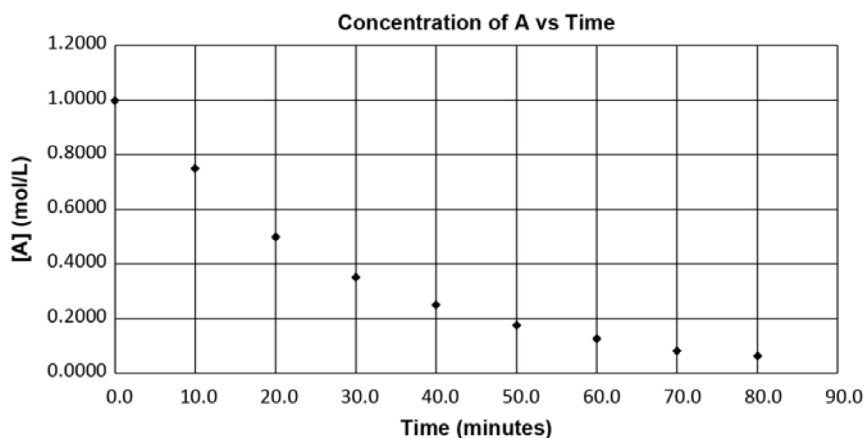
EXAMPLE:

A student conducts an experiment to measure the change in concentration of reactant A over time. The overall reaction for the system studied is:

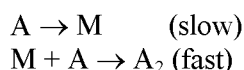


(a)

time (minutes)	[A] (molarity)
0.0	1.0000
10.0	0.7500
20.0	0.5000
30.0	0.3500
40.0	0.2500
50.0	0.1750
60.0	0.1250
70.0	0.0800
80.0	0.0625



- (i) The student states that the reaction is first order. What evidence from the data or graph can be used to support their statement?
- (ii) Calculate the rate constant for this reaction. Include units.
- (b) The mechanism shown is proposed for the reaction being studied.



Validate that the proposed mechanism supports the experimental rate law and the overall reaction.

- (c) What is the concentration of A remaining after 93 minutes?

ANS:

- (a)(i) Analysis of the data shows that the half-life is consistently 20 minutes, as the concentration is decreased by half every 20 minutes. Only first order reactions have half-lives that are independent of concentration.



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- (a)(ii) The easiest way to determine the k for this reaction is with the use of the first order half-life equation.

$$t_{\frac{1}{2}} = \frac{0.693}{k} \text{ and } k = \frac{0.693}{t_{\frac{1}{2}}}$$

$$k = \frac{0.693}{20 \text{ min}}$$

$$k = 0.035 \text{ min}^{-1}$$

- (b) The first reason why the mechanism is valid is that the rate law for the slow step (rate-determining step) matches the experimental rate law ($\text{Rate} = k[\text{A}]$). To determine the rate law for the slow step, use the coefficients (molecularity) for the reactants from the slow step. Those represent the order of reactants in the rate law.

The second reason why the mechanism is valid is that the steps shown, when added, will equal the overall reaction shown at the start of the problem. M is an intermediate that will cancel out, and the two A molecules can be combined to make 2 A, resulting in the matching overall reaction.

- (c)

$$\ln[\text{A}]_t - \ln[\text{A}]_0 = -kt$$

$$\ln[\text{A}]_{93} - \ln[1.0000 \text{ M}]_0 = -0.0347 \text{ min}^{-1}(93 \text{ min})$$

$$\ln[\text{A}]_{93} - 0 = -3.23$$

$$e^{-3.23} = 0.040 \text{ M}$$

DIF: Easy | Moderate | Difficult

OBJ: AP Learning Objective: TRA-3.C Identify the rate law expression of a chemical reaction using data that show how the concentrations of reaction species change over time. | AP Learning Objective: TRA-3.B Represent experimental data with a consistent rate law expression. | AP Learning Objective: TRA-5.A Identify the components of a reaction mechanism. | AP Learning Objective: TRA-5.B Identify the rate law for a reaction from a mechanism in which the first step is rate limiting.

TOP: AP Topic: 5.3 | AP Topic: 5.2 | AP Topic: 5.7 | AP Topic: 5.8

KEY: AP

MSC: AP Practice: 5.B | AP Practice: 5.C



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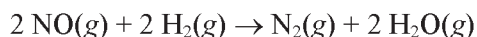
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19 AP® LONG FREE-RESPONSE QUESTIONS

1 AP® long free-response question is included per chapter.

EXAMPLE:

Nitrogen monoxide gas (NO, also known as nitric oxide), reacts with hydrogen gas, according to the following equation.



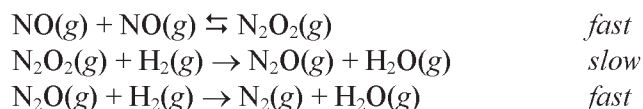
- (a) If 10.0 g H_2 reacts with excess NO according to the equation shown, calculate the theoretical yield, in grams, of N_2 that would be produced in the reaction.
- (b) A student claims the reaction shown is a redox reaction. Provide evidence to support this claim by identifying the element that is oxidized and the element that is reduced in this reaction.

Experiments were performed to study the kinetics of the reaction. All three trials were performed at the same initial temperature. The data are shown in the table.

Trial	Initial Concentration of NO(g) (M)	Initial Concentration of H ₂ (g) (M)	Initial Rate of Formation of N ₂ (g) (M s ⁻¹)
1	0.0500	0.0500	0.0104
2	0.100	0.0500	0.0417
3	0.100	0.100	0.0833

- (c) In terms of the collisions between the reactant molecules, explain why the initial rate in Trial 2 is greater than the initial rate in Trial 1. Base your answer on what occurs at the particle level.
- (d) Based on the data in the table, determine the order with respect to NO(g) and the order with respect to H₂(g). Justify your answers.
- (e) Based on your answer to part (d), write the overall rate law for the reaction.
- (f) Calculate the value of the rate constant, k , for the reaction. Include units with your answer.

The following mechanism is proposed for the reaction.



- (g) Is the rate law described by the proposed mechanism consistent with the rate law you wrote in part (e)? Justify your answer.

Substance	Standard Enthalpy of Formation, ΔH°_f (kJ/mol)
H ₂ (g)	0
NO(g)	90.3
N ₂ (g)	0
H ₂ O(g)	-241.8



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- (h) Using the table of standard enthalpies of formation, calculate the value of ΔH° for the reaction $2 \text{H}_2(\text{g}) + 2 \text{NO}(\text{g}) \rightarrow \text{N}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$.
- (i) The entropy change for this reaction, ΔS° , is $-95.5 \text{ J}/(\text{mol} \cdot \text{K})$ at 298 K. Using particle-level reasoning, explain why ΔS° is negative for this reaction.

ANS:

- (a) The theoretical yield of $\text{N}_2(\text{g})$ is calculated as follows.

$$10.0 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{1 \text{ mol N}_2}{2 \text{ mol H}_2} \times \frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} = 69.5 \text{ g N}_2$$

- (b) The oxidation numbers of the atoms in the reactants and the products are the following.
 Reactants: H = zero, N = +2, O = -2 Products: H = +1, N = zero, O = -2
 In this reaction, H is oxidized (from zero to +1) and N is reduced (from +2 to zero).
- (c) Comparing the initial concentration of the reactants between trials 1 and 2, the concentration of $\text{NO}(\text{g})$ is doubled in trial 2. As reactant concentration increases, the frequency of collisions between reactant molecules increases. This leads to a greater number of effective collisions and a faster reaction rate.
- (d) Comparing trials 1 and 2, NO is doubled while H_2 remains constant. As a result, the initial rate is quadrupled (i.e., $\frac{0.0417}{0.0104} \approx 4$). Therefore, the reaction is second order with respect to $\text{NO}(\text{g})$.
 Comparing trials 2 and 3, H_2 is doubled while NO remains constant. As a result, the initial rate is doubled (i.e., $\frac{0.0833}{0.0417} \approx 2$). Therefore, the reaction is first order with respect to $\text{H}_2(\text{g})$.

(e) $\text{Rate} = k[\text{NO}]^2[\text{H}_2]$

- (f) Using the data from trial 1:

$$\begin{aligned} \text{Rate} &= k[\text{NO}_2]^2[\text{H}_2] \\ 0.0104 \text{ M s}^{-1} &= k(0.0500 \text{ M})^2(0.0500 \text{ M}) \\ k &= \frac{0.0104 \text{ M s}^{-1}}{(0.0500 \text{ M})^2(0.0500 \text{ M})} = 83.2 \text{ M}^{-2} \text{ s}^{-1} \end{aligned}$$

Using the data from trial 2:

$$\begin{aligned} \text{Rate} &= k[\text{NO}_2]^2[\text{H}_2] \\ 0.417 \text{ M s}^{-1} &= k(0.100 \text{ M})^2(0.0500 \text{ M}) \\ k &= \frac{0.417 \text{ M s}^{-1}}{(0.100 \text{ M})^2(0.0500 \text{ M})} = 83.4 \text{ M}^{-2} \text{ s}^{-1} \end{aligned}$$

Using the data from trial 3:

$$\begin{aligned} \text{Rate} &= k[\text{NO}_2]^2[\text{H}_2] \\ 0.0833 \text{ M s}^{-1} &= k(0.100 \text{ M})^2(0.100 \text{ M}) \\ k &= \frac{0.0833 \text{ M s}^{-1}}{(0.100 \text{ M})^2(0.100 \text{ M})} = 83.3 \text{ M}^{-2} \text{ s}^{-1} \end{aligned}$$



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- (g) The rate law described by the proposed mechanism is consistent with the rate law written in part (e). Step 2 is the slow step, so it is the rate-determining step of the mechanism. The rate law of this elementary step is $\text{Rate} = k[\text{N}_2\text{O}_2][\text{H}_2]$. Because N_2O_2 is an intermediate, it cannot appear in the rate law of the overall reaction.

In step 1, the rate of the forward reaction is equal to the rate of the reverse reaction:

$$k_1[\text{NO}]^2 = k_{-1}[\text{N}_2\text{O}_2]$$

Now we can solve for N_2O_2 in terms of NO_2 and then substitute this expression into the rate law of step 2.

$$[\text{N}_2\text{O}_2] = \frac{k_1}{k_{-1}}[\text{NO}]^2$$

$$\text{Rate} = k_2[\text{N}_2\text{O}_2][\text{H}_2]$$

$$\text{Rate} = \frac{k_2 k_1}{k_{-1}}[\text{NO}]^2[\text{H}_2]$$

$$\text{Rate} = k[\text{NO}]^2[\text{H}_2]$$

This rate law is consistent with the rate law written in part (e).

- (h) The value of ΔH° for the reaction is calculated as follows.

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta H^\circ = [(0) + 2(-241.8 \text{ kJ/mol})] - [2(0) + 2(90.3 \text{ kJ/mol})]$$

$$\Delta H^\circ = -664.2 \text{ kJ/mol}_{\text{rxn}}$$

- (i) In this chemical equation, there are 4 moles of gaseous reactants and 3 moles of gaseous products. The sign of ΔS° is negative because the total number of moles of gaseous products is less than the total number of moles of gaseous reactants. This is a reduction in the number of possible arrangements of the particles, resulting in a decrease in the overall entropy for the reaction.

DIF: Moderate | Easy

OBJ: AP Learning Objective: TRA-4.B Explain the relationship between the rate of an elementary reaction and the frequency, energy, and orientation of molecular collisions. | AP Learning Objective: TRA-3.B Represent experimental data with a consistent rate law expression. | AP Learning Objective: TRA-5.C Identify the rate law for a reaction from a mechanism in which the first step is not rate limiting.

TOP: AP Topic: 4.5 | AP Topic: 4.7 | AP Topic: 5.5 | AP Topic: 5.2 | AP Topic: 5.9 | AP Topic: 6.8 | AP Topic: 9.1 KEY: AP

MSC: AP Practice: 5.F | AP Practice: 6.C | AP Practice: 6.F | AP Practice: 5.A | AP Practice: 5.C | AP Practice: 5.B | AP Practice: 6.C



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