

# Part I

## Strategies for the AP Exam

The best way to start with an AP examination is to figure out what the exam is all about. If you have taken the exam before, you know what to expect. If you are taking it for the first time, you need to know what to expect. The best way to start is to read the book. The book is divided into two parts. Part I is for students who are taking the exam for the first time. Part II is for students who are taking the exam for the second time. The book is divided into two parts. Part I is for students who are taking the exam for the first time. Part II is for students who are taking the exam for the second time. The book is divided into two parts. Part I is for students who are taking the exam for the first time. Part II is for students who are taking the exam for the second time.

### WHAT'S IN THIS BOOK

This book is divided into two parts. Part I is for students who are taking the exam for the first time. Part II is for students who are taking the exam for the second time. The book is divided into two parts. Part I is for students who are taking the exam for the first time. Part II is for students who are taking the exam for the second time. The book is divided into two parts. Part I is for students who are taking the exam for the first time. Part II is for students who are taking the exam for the second time.

Part I is divided into 11 chapters—again following the College Board's outline. These chapters are not a substitute for your textbook and they shouldn't be. They are meant to be used as a supplement. At the end of each chapter, you will find 15 multiple-choice questions and two free-response questions based on the material in that chapter. At the end of each chapter, you will find 15 multiple-choice questions and two free-response questions based on the material in that chapter. At the end of each chapter, you will find 15 multiple-choice questions and two free-response questions based on the material in that chapter.

AP Chemistry Practice Exam

# PREPARING FOR THE AP<sup>®</sup> CHEMISTRY EXAMINATION

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Advanced Placement can be exhilarating. Whether you are taking an AP course at your school or you are working on AP independently, the stage is set for a great intellectual experience.

But sometime after New Year's Day, when the examination begins to loom on a very real horizon, Advanced Placement can seem downright intimidating—in fact, offered the opportunity to take the examination just to see what it's like, even adults long out of high school refuse. If you dread taking the test, you are in good company.

The best way to deal with an AP examination is to master it, not let it master you. If you can think of these examinations as a way to show off how much chemistry you know, you have a leg up. Attitude and confidence *do* help. If you are not one of those students, there is still a lot you can do to sideline your anxiety. This book is designed to put you on a fast track. Focused review and practice time will help you master the examination so that you can walk in with confidence and score a 5.

## WHAT'S IN THIS BOOK

This book is keyed to *Chemistry* by Steven and Susan Zumdahl, 8th and 9th editions, but because it follows the College Board Concept Outline, it is compatible with all textbooks. It is divided into three sections. Part I offers suggestions for getting yourself ready, from signing up to take the test and sharpening your pencils to organizing a free-response essay. This is followed by the list of AP Chemistry learning objectives from the College Board curriculum. At the end of Part I, you will find a Diagnostic Test that has all of the elements of the AP Chemistry Examination.

Part II is made up of 13 chapters—again following the College Board Concept Outline. These chapters are not a substitute for your textbook and class discussion; they simply review the AP chemistry course. At the end of each chapter, you will find 15 multiple-choice questions and two free-response questions based on the material in that chapter. At the end of each answer, you will find page references directing you to the discussion on that particular point in *Chemistry* and the AP Chemistry learning objective number covering that point.

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Part III has two complete AP Chemistry examinations. At the end of each test, you will find the answers, explanations, and references to *Chemistry* and the AP Chemistry learning objective for the multiple-choice and the free-response questions.

## WHAT'S IN THE CHEMISTRY TEXTBOOK THAT WILL HELP YOUR PREPARATION

As you work your way through the textbook there are some features that will assist you in getting the most out of the textbook:

- Make use of the Conceptual Problem Solving Method introduced in Chapter 3 in which you break the problem down into three parts: (1) Where are we going? (2) How do we get there? and (3) Reality Check.
- Make certain you carefully study the Sample Exercises with solutions in each chapter. These will often give you a guide to help develop your problem solving skills.
- Read the For Review section that highlights the material presented in the chapter as a means to double check your understanding before you proceed.
- Answer the Review Questions at the end of each chapter as a measure of your understanding. This will give you instant feedback as to how much of the chapter you may need to study again.
- Perform the multiple-choice AP assessment items at the end of each chapter. They are tied to the Learning Objectives which are part of the AP format. Practice doing computations for these questions without a calculator. You are not allowed to use a calculator in the multiple-choice section of the test, so you need to be good at estimating.
- Working in small groups, answer the Active Learning Questions at the end of each chapter. This is an excellent opportunity to gauge your level of understanding and to receive assistance from your classmates.
- Do as many Challenge and Marathon Problems as you can. These problems are intended to incorporate many different concepts into the solution. Problems like this are particularly good practice for the free-response portion of the AP exam.

Being successful in chemistry usually involves careful reading of the textbook and working as many different types of problems as possible. The successful student is not one who memorizes a bunch of facts but one who is able to synthesize and analyze the problem. Chemistry is like a lot of subjects in that an understanding of new material often depends on an understanding of previous concepts.

## SETTING UP A REVIEW SCHEDULE

If you have been doing your homework steadily and keeping up with the course work, you are in good shape. Organize your notes, homework, and handouts from class by topic. Reference these materials as well as your textbook and this study guide when you have difficulty in a specific section. But even if you have done all that—or if

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it is too late to do all that—there are some more ways to get it all together.

To begin, read Part I of this book. You will be much more comfortable going into the exam if you understand how the exam questions are designed and how best to approach them. Then take the Diagnostic Test and see where you are right now.

Take out a calendar and set up a schedule for yourself. If you begin studying early, you can chip away at the review chapters in Part II. You will be surprised—and pleased—by how much material you can cover in half an hour a day of study for a month or so before the test. Look carefully at the sections of the Diagnostic Test; if you missed a number of questions in one particular area, allow more time for the chapters that cover that area of the course. The practice tests in Part III will give you more experience with different kinds of multiple-choice questions and the wide range of free-response questions.

If time is short, skip reading the review chapters (although you might read through the chapter subheadings) and work on the multiple-choice and free-response questions at the end of each review. This will give you a good idea of your understanding of that particular topic. Then take the tests in Part III.

If time is *really* short, go straight from Part I to Part III. Taking practice tests over and over again is the fastest, most practical way to prepare. You cannot study chemistry by reading it like a novel. You must actively do problems to gain understanding and excel in your performance. Athletes do not perform well just by reading books about their sport or by watching others. They must get up and practice. So, you too, just like athletes, must practice, practice, practice if you want to do your best!

### BEFORE THE EXAMINATION

By February, long before the exam, you need to make sure that you are registered to take the test. Many schools take care of the paperwork and handle the fees for their AP students, but check with your teacher or the AP coordinator to make sure that you are on the list. This is especially important if you have a documented disability and need test accommodations. If you are studying AP independently, call AP Services at the College Board for the name of the local AP coordinator, who will help you through the registration process.

The evening before the exam is not a great time for partying. Nor is it a great time for cramming. If you like, look over class notes or drift through your textbook, concentrating on the broad outlines, not the small details, of the course. You might also want to skim through this book and read the AP tips.

The evening before the exam is a great time to get your things together for the next day. Sharpen a fistful of no. 2 pencils with good erasers; bring a scientific calculator with fresh batteries. Certain types of calculators are not allowed, so be sure to verify with your teacher or the College Board that your model is acceptable. For example, you cannot use a calculator with a typewriter-style keyboard or a cell phone. Cell phones are not even allowed in the testing room, so you will need a watch and be certain to turn off the alarm if it has one. Bring a piece of fruit or a power bar and a bottle of water for the

break. Make sure you have your Social Security number and whatever photo identification and admission ticket are required. Then relax. And get a good night's sleep.

On the day of the examination, plan to arrive early. It is wise not to skip breakfast—studies show that students who eat a hot breakfast before testing get higher grades. Be careful not to drink a lot of liquids, necessitating a trip to the bathroom during the exam. Breakfast will give you the energy you need to power you through the exam—and more. You will spend some time waiting while everyone is seated in the right room for the right exam before the test has even begun. With a short break between Section I and Section II, the AP Chemistry exam lasts for more than two and a half hours. So be prepared for a long morning. You do not want to be distracted by a growling stomach or hunger pangs.

Be sure to wear comfortable clothes, taking along a sweater in case the heating or air-conditioning is erratic. Be sure, too, to wear clothes you like—everyone performs better when they think they look better—and by all means wear your lucky socks.

You have been on the fast track. Now go get a 5!

Section	Number of questions	Time to complete	Number of choices	Calculator	Weighting
I	60	90 minutes	5	Yes	30%
II	1	1 hour	1	No	70%

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I	60	90 minutes	5	Yes	30%
II	1	1 hour	1	No	70%

# TAKING THE AP<sup>®</sup> CHEMISTRY EXAMINATION

The AP Chemistry curriculum emphasizes the importance of inquiry and reasoning in science, and requires depth of understanding more than the memorization of facts. On the exam you may be asked to do a range of tasks, such as interpreting lab data or making predictions based on your knowledge of chemical principles.

Some content in your text will not be tested in the examination, although your teacher may choose to include that material in your course.

The AP Chemistry exam will consist of two sections. The sections are described below.

**Section I:** The main features of the multiple-choice sections are given in the following table.

	<b>Format</b>
Number of questions	60
Time to complete	90 minutes
Number of choices	4
Calculator	No
Weighting	50%

The questions will generally refer to “real world” situations and may contain several parts. Some sample questions are given in the “Strategies for the Multiple-Choice Section” that follows.

**Section II:** After Section I is collected, you will have a short break. The main features of the free-response sections are given in the following table.

	<b>Format</b>
Number of questions	7 total. The questions are made up of 3 long questions (parts a–d or parts a–e) and 4 short questions (parts a–b).

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	<b>Format</b>
Point value	This section is worth up to 46 points. Each long question is worth up to 10 points. Each short question is worth up to 4 points.
Types of long and short questions (at least one question from each type)	Lab I Lab II Representation I Representation II Quantitative
Time to complete	90 minutes
Calculator	Yes
Weighting	50%

In this section, you will be asked to:

- Design an experiment to test a hypothesis, given a specific set of laboratory equipment (Lab I)
- Interpret and draw conclusions from data (Lab II)
- Use models to describe physical phenomena (Representations I and II)
- Follow a logical or analytical pathway to solve a problem (Quantitative)

When answering these questions, you will need to be able to explain, say, what is happening at the atomic or molecular level when a substance changes from a solid, to a liquid, to a gas; but in addition, you may be asked to draw diagrams of the three phases to demonstrate that you can make models of systems.

The focus on laboratory work as a tool of scientific inquiry means that you will need to practice writing laboratory procedures, understand the correct use of laboratory equipment to gather data, and incorporate safe practices. Some example questions are given in the “Strategies for Answering Free-Response Questions” that follows.

## STRATEGIES FOR THE MULTIPLE-CHOICE SECTION

Each question in the multiple-choice section is worth one point, and since there is no penalty for guessing, you want to make the most educated guess! Here are some rules of thumb to help you:

- **Read the question carefully** Pressured for time, many students make the mistake of reading the questions too quickly or merely skimming them. By reading a question carefully, you may already have some idea about the correct answer. You can then look for it in the responses.

- **Eliminate any answer you know is wrong** You can write on the multiple-choice questions in the test book. As you read through the responses, draw a line through any answer you know is wrong.
- **Read all of the possible answers, then choose the most accurate response** AP examinations are written to test your precise knowledge of a subject. Some of the responses may be partially correct but there will only be one response that is completely true.
- **Avoid absolute responses** These answers often include the words “always” or “never.” For example, the statement “all chlorides are always soluble in water” is incorrect because compounds such as silver chloride and lead (II) chloride are insoluble in water.
- **Mark tough questions** If you are hung up on a question, make an educated guess, but mark it in the margin of the question book and come back to review it later if you have time.

## TYPES OF MULTIPLE-CHOICE QUESTIONS

The exam will include short and long questions. The latter will ask several questions about the same set of data. Here are some suggestions for approaching each of the various kinds of multiple-choice questions:

**CLASSIC/BEST ANSWER QUESTIONS** This is the most common type of multiple-choice question. It simply requires you to read the question and select the most correct answer. The example below is a short question.

1. Given the following reduction potentials, which metal would be best for a pipeline carrying dilute hydrochloric acid?

Reaction	$E^\circ, \text{V}$
$\text{Ti}^{3+} + 3\text{e}^- \rightarrow \text{Ti}$	+0.72
$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.76
$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	+0.34
$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	-0.44

- (A) Ti
- (B) Zn
- (C) Cu
- (D) Fe

**ANSWER: A.** You are looking for the metal that is least likely to be oxidized by the protons in a solution of the strong acid. The higher (most positive) the reduction potential, the less likely the metal will be oxidized. Eliminate (B) and (D) because their reduction potentials are negative, indicating that they are more likely to be oxidized. Of the two remaining, (A) has the largest positive value. (Learning Objective 3.12: The student can make qualitative or quantitative predictions about galvanic or electrolytic reactions based on half-cell reactions and potentials and/ or Faraday’s laws.)



**NONCALCULATOR COMPUTATIONS** These questions require computation without the use of a calculator. Simple mathematics or the choice of the correct algebraic setup will be involved.

1. A weak acid, HA, has a  $K_a$  value of  $1.0 \times 10^{-6}$ . A student adds 10.0 mL of 0.200 M NaOH to 40.0 mL of 0.100 M HA and the two react completely. Calculate the final pH.
  - (A) 2.0
  - (B) 6.0
  - (C) 7.0
  - (D) 8.0

**ANSWER: B.** To answer this question, you must consider at what point in the titration the question is referring to. Because the base, NaOH, is twice as concentrated as the acid, HA, it will take half as much NaOH to reach the equivalence point, or 20.0 mL. We are not at the equivalence point. Half the amount of the NaOH, 10.0 mL, needed to reach the equivalence point has been added. Halfway to the equivalence point, the pH equals the  $pK_a$  which equals  $-\log(K_a)$  or 6.0. (Learning Objective 6.13: The student can interpret titration data for monoprotic or polyprotic acids involving titration of a weak or strong acid by a strong base (or a weak or strong base by a strong acid) to determine the concentration of the titrant and the  $pK_a$  for a weak acid, or the  $pK_b$  for a weak base.)

2.  $2\text{C}(s) + \text{O}_2(g) \rightarrow 2\text{CO}(g)$

Into a 3.0-L container at 25°C are placed 1.2 g of carbon (graphite), and 3.2 g of oxygen gas.

If the carbon and the oxygen react completely to form CO(g), what will be the final pressure in the container at 25°C?

- (A)  $\frac{0.05(0.082)(298)}{3.0}$  atm
- (B)  $\frac{0.10(0.082)(25)}{3.0}$  atm
- (C)  $\frac{0.15(0.082)(298)}{3.0}$  atm
- (D)  $\frac{0.10(0.082)(298)}{3.0}$  atm

**ANSWER: C.** To answer this question, 1.2 g of carbon is 0.10 mol of carbon (1 mole of carbon equals 12 g). Similarly, 3.2 g of  $\text{O}_2$  is 0.10 mol of  $\text{O}_2$ . The limiting reactant is carbon producing 0.10 mol of CO.  $\text{O}_2$  is present in excess. (0.10 mol of CO requires 0.05 mol of  $\text{O}_2$  leaving 0.05 mol of  $\text{O}_2$ .) When the reaction is complete, there is 0.10 mol CO + 0.05 mol  $\text{O}_2$  left which equals 0.15 mol total. (Learning Objective 2.6: The student can apply mathematical relationships or estimation to determine macroscopic variables for ideal gases. Learning Objective 3.4: The student is able to relate quantities (measured mass of substances, volumes of solutions, or volumes and pressures of gases) to identify stoichiometric relationships for a reaction, including situations involving limiting reactants and situations in which the reaction has not gone to completion.)

## Long Question

Questions 1 and 2 refer to the following data.

A student performing an experiment to determine the relative intermolecular forces of various liquids immerses a thermocouple covered with filter paper in a specific liquid, and monitors the resulting temperature change over a 3-minute interval. The data she obtains are:

Liquid	$\Delta T$ ( $^{\circ}\text{C}$ )
Ethanol	8.0
1-Propanol	7.5
2-Propanol	8.0
Water	3.5

- Based on these data, which substance has the strongest intermolecular forces?
  - Ethanol
  - 1-Propanol
  - 2-Propanol
  - Water

**ANSWER: D.** Substances with strong intermolecular forces do not readily change phase from liquid to gas. Temperature change, specifically cooling, is related to the amount of substance that evaporates, so a low temperature change indicates less evaporation and higher intermolecular forces. Water exhibits the smallest temperature change so water must have the strongest intermolecular forces. (Learning Objective 2.13: The student is able to describe the relationships between the structural features of polar molecules and the forces of attraction between the particles.)

- Why does 2-propanol have a greater temperature change than 1-propanol?
  - 1-Propanol has more London dispersion forces because it has a longer carbon chain.
  - 2-Propanol has more London dispersion forces because it has a longer carbon chain.
  - The hydrogen bonds are stronger in 1-propanol than in 2-propanol.
  - The hydrogen bonds are stronger in 2-propanol than in 1-propanol.

**ANSWER: A.** Both substances are able to form the same number of hydrogen bonds per molecule, so this does not explain the difference in temperature change and eliminates choices (C) and (D). 1-Propanol has a straight three-carbon chain so it has more contact points available for London dispersion forces than the branched chain of 2-propanol. (Learning Objective 2.16: The student is able to explain the properties (phase, vapor pressure, viscosity, etc.) of small and large molecular compounds in terms of the strengths and types of intermolecular forces.)

## STRATEGIES FOR THE FREE-RESPONSE SECTION

Section II of the AP exam comes with a periodic table and a table of equations and constants. The needed portions of the standard reduction potentials,  $E^{\circ}_{\text{red}}$ , will be provided within the questions.

- Scan all of the questions in the section you are working in and mark those that you know you can answer correctly. Do these problems first.
- Show all of your work and use units where requested. Partial credit will be awarded for problems if the correct work is shown but the answer is not present or is incorrect. In problems involving calculations, circle your final answer.
- Cross out incorrect answers with an “X” rather than spending time erasing.
- Be clear, neat, and organized in your work. If a grader cannot clearly understand your work, you may not receive full credit.
- Three of the questions will have several parts. Attempt to solve each part. Even if your answer to the first part is incorrect, you still may be awarded points for the remaining parts of the question if the work is correct for those parts.
- Units are important in your answer. Keeping track of your units throughout calculations, and performing unit cancellation where possible, will help guide you to your answer. Points will be deducted for missing or incorrect units in the answer. However, if the question does not specify the requirement of units, points will not be deducted for missing units, although they can be deducted for incorrect units.
- You do not need to work the questions in order. Make sure you put the number of the question in the corner of each page of your essay booklet. In addition, questions are sometimes broken into parts, such as (a) and (b). When this is the case, label each part of your response.

## STRATEGIES FOR ANSWERING FREE-RESPONSE QUESTIONS

Free-response questions will ask you to explain, compare, and predict. Minor calculations, showing mathematical relationships, or drawing graphs or structures, may also be involved. Usually these are not traditional essay questions. Most free-response questions do not require an introduction or conclusion. Most do not even require a thesis. Many of these questions may be written in a bulleted or short-answer format. Although this may sound easier than writing a traditional essay, it is important that you know the material very well because these are targeted questions. Examination readers want specifics. They are looking for accurate information presented in clear, concise prose. You cannot mask vague information with elegant prose. The use of appropriate vocabulary terms is strongly suggested as this may make the difference between the awarding of a point or not.

To be successful in writing free-response answers for the AP Chemistry exam, be sure to get straight to the answer and use key terms in your explanations. Sometimes, if you ramble on and on, you

might accidentally state an incorrect fact. Points will be deducted for incorrect or extraneous information.

Techniques to write free-response questions include chart format, bullet format, and outline format. None of these styles requires that you write complete sentences. In each of these styles, restate the question in simple terms, using your own words. Restate each part (a, b, and so on) separately, not together. In your restatement and response, underline key words or concepts.

- **Fill in a chart to answer the question.** This style is helpful in answering questions about electronic and molecular structure. This method is also used in the example on identification of the set of four quantum numbers for each electron in an element in this chapter.
- **In bullet or number format, make a list,** using a bullet (■) or number for each new concept. Leave room between concepts because you may want to come back later to fill them in.
- **Outline format** is more traditional, using Roman numerals, letters, etc. This takes more time to organize ideas, but it does show progression of ideas in a logical sequence. As in the bulleted format, leave room between concepts because you may want to come back later to fill them in.
- Lastly, **the freestyle method,** but this sometimes results in rambling and incomplete answers. Also, writing in paragraphs does not allow room for additional ideas to be added.

## TYPES OF FREE-RESPONSE QUESTIONS

The free-response questions may be long or short. The long questions typically include four or five parts and are worth up to 10 points. The short questions include one or two parts and are worth up to 4 points. Samples and descriptions of the main categories of questions follow.

**SAMPLE PROBLEM, LAB I:** In this type of question, you will be asked to demonstrate your ability to plan and write a laboratory procedure based on your knowledge of chemical principles. The graders will also be looking for your knowledge of safe practices. These can be long or short questions, but will often be long questions, as in the example below.

You are given a set of four clear liquids. One is ethanol, one is a dilute solution of acetic acid, one is water, and one is a dilute solution of sodium hydroxide. Design an experiment to identify the liquids. Describe the steps, the data you would collect, and how the data support the identifications. Laboratory equipment for your experiment should be taken from the list below. (You may not need all of the equipment.)

50-mL beakers	fume hood
hot plate	stirring rods
0.1 M NaHCO <sub>3</sub>	phenolphthalein indicator
conductivity meter	safety glasses
thermometer	beaker tongs

**Answer:**

1. Work under a fume hood, because several of the liquids are volatile.
2. Wear safety glasses to protect your eyes from the acid, base, and alcohol.
3. Put approximately 5 mL of each liquid in each of 4 50-mL beakers.
4. First test: Of the liquids, only sodium hydroxide will turn pink in the presence of phenolphthalein indicator. Add a drop of indicator to each solution and stir gently. The one that turns pink is sodium hydroxide solution. Clean off the stirring rods with distilled water so they will not cause contamination in subsequent tests. (Learning Objective 6.15: The student can identify a given solution as containing a mixture of strong acids and/or bases and calculate or estimate the pH—and concentrations of all chemical species—in the resulting solution.)
5. Second test: Of the remaining liquids, only acetic acid will react with sodium hydrogen carbonate to release carbon dioxide gas. The phenolphthalein indicator will not interfere with the test, so the same beakers can be used. Add approximately 5 mL of sodium hydrogen carbonate solution to the three remaining solutions and stir gently. The one that bubbles is the acetic acid being neutralized by the weak base to form water, carbon dioxide gas, and aqueous sodium acetate. (Learning Objective 6.16: The student can identify a given solution as being the solution of a monoprotic weak acid or base (including salts in which one ion is a weak acid or base), calculate the pH and concentration of all species in the solution, and/or infer the relative strengths of the weak acids or bases from given equilibrium concentrations.)
6. Third test: Of the remaining liquids, ethanol has a much lower boiling temperature than water because it has weaker intermolecular forces (it can only make one hydrogen bond per molecule, and water can make two). Put approximately 10 mL of each in two separate 50-mL beakers and put on a hot plate set at medium temperature. Check the temperatures of the liquids when they boil by placing a thermometer in the liquid, but not touching the bottom or sides of the beaker. The water should boil at around 90–100°C; the ethanol will boil at a much lower temperature. Use the beaker tongs to remove the beakers from the hot plate as soon as you are done making measurements. (Learning Objective 2.16: The student is able to explain the properties (phase, vapor pressure, viscosity, etc.) of small and large molecular compounds in terms of the strengths and types of intermolecular forces.)
7. Dispose of all liquids in the designated liquid waste containers in your lab.

Solving this problem also involves the use of Science Practice 4: The student can plan and implement data collection strategies in relation to a particular scientific question. [Note: Data can be collected from many different sources, e.g., investigations, scientific observations, the findings of others, historic reconstruction, and/or archived data.] The

problem also involves the use of Science Practice 5: The student can perform data analysis and evaluation of evidence.

**PROBLEM, LAB II:** Here you will be given data and asked to draw conclusions based on these data. Your answers must demonstrate your knowledge of the specific chemical principles that relate to the observations. Using proper vocabulary is critical. These questions may be long or short. The example is a long question.

You are given a set of four vials of clear liquids, designated A, B, C, and D. One is ethanol, one is a dilute solution of acetic acid, one is mineral oil, and one is a dilute solution of sodium hydroxide. In a series of experiments, you collect the following data:

- Only A and B are strongly conductive.
- Only B turns pink when phenolphthalein indicator is added.
- When equal portions of each liquid are placed on thermocouples and the amount of cooling tracked over a 5-minute period, C shows the greatest change.
- When small portions of the liquids are mixed with water, all form homogeneous solutions except D, which forms a separate layer on the top.

Use this information to identify the contents of each vial. Explain clearly how you used the data to come to your conclusions.

**Answer:**

1. Compounds that ionize when dissolved in water form conductive solutions. Of the four liquids, only the sodium hydroxide solution (a base) and the acetic acid solution (an acid) consist of ions. A and B must be the acid and base. Since B turns pink when phenolphthalein is added, and phenolphthalein is only pink in basic solutions, B must be sodium hydroxide and A must be acetic acid. (Learning Objective 2.15: The student is able to explain observations regarding the solubility of ionic solids and molecules in water and other solvents on the basis of particle views that include intermolecular interactions and entropic effects. Learning Objective 2.16: The student is able to explain the properties (phase, vapor pressure, viscosity, etc.) of small and large molecular compounds in terms of the strengths and types of intermolecular forces.)
2. Ethanol is completely soluble in water because it can form hydrogen bonds with water. Mineral oil is nonpolar and will not dissolve in water. D must be mineral oil because it forms a layer on the top of water. (Learning Objective 2.15: The student is able to explain observations regarding the solubility of ionic solids and molecules in water and other solvents on the basis of particle views that include intermolecular interactions and entropic effects.)
3. Although ethanol forms hydrogen bonds with itself, it can only make one hydrogen bond per molecule and water can make two. With fewer intermolecular forces per molecule, ethanol will be more volatile than water and the aqueous acid and base solutions. Mineral oil is nonpolar but as a linear molecule has

many contact points for London dispersion forces; it is also relatively nonvolatile. Ethanol will evaporate the most in a set period of time and therefore cool the thermocouple the most. C must be ethanol. (Learning Objective 2.16: The student is able to explain the properties (phase, vapor pressure, viscosity, etc.) of small and large molecular compounds in terms of the strengths and types of intermolecular forces.)

Solving this problem also demonstrates Science Practice 5: The student can perform data analysis and evaluation of evidence.

Scientists, chemists in particular, create models of atomic/molecular behavior that explain what is observed at “large” scale. In Representation problems, you are asked to talk about the strengths and weaknesses of key models. There are two types of Representation questions.

**SAMPLE PROBLEM, REPRESENTATION I:** For Representation I questions you are asked to discuss how a model of particles behave on the microscopic (particulate) level and explain macroscopic observations. You will need to understand the strengths and the weaknesses of various models. These can be long or short questions. The example below is a short question.

In the kinetic-molecular theory (KMT), gas atoms/molecules are depicted as particles with no appreciable volume, in constant random motion, that collide elastically with one another and do not interact with each other. What aspect(s) of this model break(s) down when a system’s temperature decreases and pressure increases?

**Answer:**

1. In KMT, the number and size of molecules are so small relative to the size of the container that their volume can be neglected, and the system volume is essentially that of the “empty” container. The main way to increase pressure is to decrease the system volume. If the system volume decreases enough, the volume of the molecules themselves is no longer negligible and must be taken into account. Also, the molecules may get close enough for intermolecular forces to come into play, so the molecules now are able to interact with each other.
2. Temperature is a measure of molecular motion, so a lower temperature indicates that the molecules are moving more slowly. When they move more slowly, they can spend enough time near each other for intermolecular forces to become significant. (Learning Objective 2.4: The student is able to use KMT and concepts of intermolecular forces to make predictions about the macroscopic properties of gases, including both ideal and nonideal behaviors.)

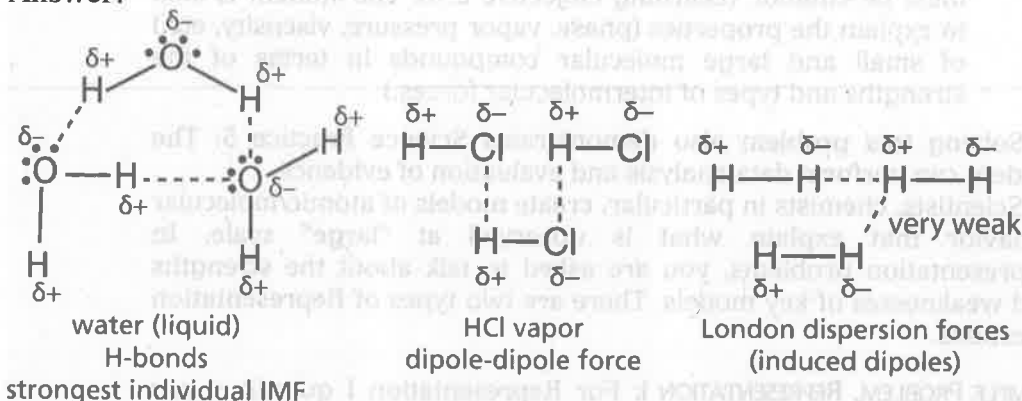
**SAMPLE PROBLEM, REPRESENTATION II:** This kind of problem is similar to Representation I, except that your answer will contain diagrams as well as explanations. These can be long or short questions. This example is a short question.

Water is a liquid at room temperature, but hydrogen chloride and hydrogen are gases under the same conditions. Draw a particulate representation of each substance, clearly indicating the intermolecular

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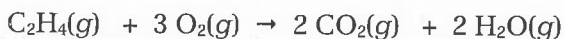
forces that account for these observations, and name the primary intermolecular force for each.

Answer:



(Learning Objective 2.1: Students can predict properties of substances based on their chemical formulas, and provide explanations of their properties based on particle views. Learning Objective 2.4: The student is able to use KMT and concepts of intermolecular forces to make predictions about the macroscopic properties of gases, including both ideal and nonideal behaviors.)

**SAMPLE PROBLEM, QUANTITATIVE:** This type of question will ask you to follow a logical, analytical pathway to solve a problem. These can be long or short questions. The example given is a long question.



Information about the substances involved in the reaction presented above is tabulated below.

Substance	$S^\circ(\text{J/mol}\cdot\text{K})$	$G_f^\circ(\text{kJ/mol})$	Bond	Bond Energy (kJ/mol)
$\text{C}_2\text{H}_4(g)$	?	68	O - H	467
$\text{O}_2(g)$	205	0	O - O	146
$\text{CO}_2(g)$	213.6	-394	O = O	495
$\text{H}_2\text{O}(g)$	189	-229	C - H	413
			C - C	347
			C = C	614
			C ≡ C	839
			C - O	358
			C = O	799



- (a) Calculate the value for the standard free energy change,  $\Delta G^\circ$ , at 25°C for the reaction. What does the sign of  $\Delta G^\circ$  indicate about the reaction?
- (b) Calculate the value for the standard enthalpy change,  $\Delta H^\circ$ , at 25°C for the reaction. What does the sign of  $\Delta H^\circ$  indicate about the reaction?
- (c) Calculate the value for the standard entropy change,  $\Delta S^\circ$ , at 25°C for the reaction. What does the sign of  $\Delta S^\circ$  indicate about the reaction?
- (d) Calculate the value for the absolute entropy of  $C_2H_4(g)$  at 25°C.

**ANSWER**

$$(A) \Delta G^\circ = 2(-394) + 2(-229) - 68 = -1314 \text{ kJ/mol}$$

(B)


Bonds Broken

$$4 \text{ C-H} \quad 4(413)$$

$$1 \text{ C=C} \quad 1(614)$$

$$3 \text{ O=O} \quad \underline{3(495)}$$

$$+3751 \text{ kJ}$$

Bonds Formed

$$4 \text{ C=O} \quad 4(799)$$

$$\underline{4 \text{ O-H}} \quad \underline{4(467)}$$

$$-5064 \text{ kJ}$$

$$\Delta H^\circ = 3751 \text{ kJ} + (-5064 \text{ kJ}) = -1313 \text{ kJ}$$

The reaction is exothermic since  $\Delta H^\circ$  is negative.

$$(C) \Delta G^\circ = \Delta H^\circ - T\Delta S^\circ; \quad \Delta S^\circ = (\Delta H^\circ - \Delta G^\circ) / T$$

$$\Delta S^\circ = -1313 - (-1314 \text{ kJ}) / 298 \text{ K} = 0.0034 \text{ kJ/K}$$

A positive sign of change in entropy indicates that entropy is increasing in the reaction.

Here, it barely changes.

$$(D) \Delta S^\circ = 2 S^\circ_{\text{CO}_2} + 2 S^\circ_{\text{H}_2\text{O}} - (S^\circ_{\text{C}_2\text{H}_4} + 3 S^\circ_{\text{O}_2})$$

$$S^\circ_{\text{C}_2\text{H}_4} = 2 S^\circ_{\text{CO}_2} + 2 S^\circ_{\text{H}_2\text{O}} - 3 S^\circ_{\text{O}_2} - \Delta S^\circ$$

$$= 2(213.6) + 2(189) - 3(205) - 3.4 = 187 \text{ J/K}$$

Learning Objective 5.8: The student is able to draw qualitative and quantitative connections between the reaction enthalpy and the energies involved in the breaking and formation of chemical bonds.

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Learning Objective 5.13: The student is able to predict whether or not a physical or chemical process is thermodynamically favored by determination of (either quantitatively or qualitatively) the signs of both  $\Delta H^\circ$  and  $\Delta S^\circ$ , and calculation or estimation of  $\Delta G^\circ$  when needed.

Learning Objective 5.14: The student is able to determine whether a chemical or physical process is thermodynamically favorable by calculating the change in standard Gibbs free energy.



Bond	Number	Energy (kJ/mol)
C-H	4	413
C-C	1	347
C-O	2	358
O-H	2	463
O=O	1	498
H-H	1	436

The reaction is exothermic since  $\Delta H$  is negative.

(c)  $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = 26 \text{ kJ} - (298 \text{ K})(-0.075 \text{ kJ/K}) = 47 \text{ kJ}$

$\Delta S^\circ = -118 \text{ J/K} - (-131 \text{ J/K}) = 0.003 \text{ kJ/K}$

A positive sign of change in entropy indicates that entropy is increasing in the reaction.

There is a small change.

(d)  $\Delta S^\circ = 5 \text{ J/K} + 5 \text{ J/K} - 10 \text{ J/K} = 0 \text{ J/K}$

$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = 26 \text{ kJ} - (298 \text{ K})(0 \text{ kJ/K}) = 26 \text{ kJ}$

Learning Objective 5.14: The student is able to determine whether a chemical or physical process is thermodynamically favorable by calculating the change in standard Gibbs free energy.

# AP CHEMISTRY

## LEARNING OBJECTIVES

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The following is the list of learning objectives from the College Board curriculum.

**LEARNING OBJECTIVE 1.1** The student can justify the observation that the ratio of the masses of the constituent elements in any pure sample of that compound is always identical on the basis of the atomic molecular theory.

**LEARNING OBJECTIVE 1.2** The student is able to select and apply mathematical routines to mass data to identify or infer the composition of pure substances and/or mixtures.

**LEARNING OBJECTIVE 1.3** The student is able to select and apply mathematical relationships to mass data in order to justify a claim regarding the identity and/or estimated purity of a substance.

**LEARNING OBJECTIVE 1.4** The student is able to connect the number of particles, moles, mass, and volume of substances to one another, both qualitatively and quantitatively.

**LEARNING OBJECTIVE 1.5** The student is able to explain the distribution of electrons in an atom or ion based upon data.

**LEARNING OBJECTIVE 1.6** The student is able to analyze data relating to electron energies for patterns and relationships.

**LEARNING OBJECTIVE 1.7** The student is able to describe the electronic structure of the atom, using PES data, ionization energy data, and/or Coulomb's law to construct explanations of how the energies of electrons within shells in atoms vary.

**LEARNING OBJECTIVE 1.8** The student is able to explain the distribution of electrons using Coulomb's law to analyze measured energies.

**LEARNING OBJECTIVE 1.9** The student is able to predict and/or justify trends in atomic properties based on location on the periodic table and/or the shell model.

**LEARNING OBJECTIVE 1.10** Students can justify with evidence the arrangement of the periodic table and can apply periodic properties to chemical reactivity.

**LEARNING OBJECTIVE 1.11** The student can analyze data, based on periodicity and the properties of binary compounds, to identify patterns and generate hypotheses related to the molecular design of compounds for which data are not supplied.

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**LEARNING OBJECTIVE 1.12** The student is able to explain why a given set of data suggests, or does not suggest, the need to refine the atomic model from a classical shell model with the quantum mechanical model.

**LEARNING OBJECTIVE 1.13** Given information about a particular model of the atom, the student is able to determine if the model is consistent with specified evidence.

**LEARNING OBJECTIVE 1.14** The student is able to use data from mass spectrometry to identify the elements and the masses of individual atoms of a specific element.

**LEARNING OBJECTIVE 1.15** The student can justify the selection of a particular type of spectroscopy to measure properties associated with vibrational or electronic motions of molecules.

**LEARNING OBJECTIVE 1.16** The student can design and/or interpret the results of an experiment regarding the absorption of light to determine the concentration of an absorbing species in a solution.

**LEARNING OBJECTIVE 1.17** The student is able to express the law of conservation of mass quantitatively and qualitatively using symbolic representations and particulate drawings.

**LEARNING OBJECTIVE 1.18** The student is able to apply conservation of atoms to the rearrangement of atoms in various processes.

**LEARNING OBJECTIVE 1.19** The student can design, and/or interpret data from, an experiment that uses gravimetric analysis to determine the concentration of an analyte in a solution.

**LEARNING OBJECTIVE 1.20** The student can design, and/or interpret data from, an experiment that uses titration to determine the concentration of an analyte in a solution.

**LEARNING OBJECTIVE 2.1** Students can predict properties of substances based on their chemical formulas, and provide explanations of their properties based on particle views.

**LEARNING OBJECTIVE 2.2** The student is able to explain the relative strengths of acids and bases based on molecular structure, interparticle forces, and solution equilibrium.

**LEARNING OBJECTIVE 2.3** The student is able to use aspects of particulate models (i.e., particle spacing, motion, and forces of attraction) to reason about observed differences between solid and liquid phases and among solid and liquid materials.

**LEARNING OBJECTIVE 2.4** The student is able to use KMT and concepts of intermolecular forces to make predictions about the macroscopic properties of gases, including both ideal and nonideal behaviors.

**LEARNING OBJECTIVE 2.5** The student is able to refine multiple representations of a sample of matter in the gas phase to accurately represent the effect of changes in macroscopic properties on the sample.

**LEARNING OBJECTIVE 2.6** The student can apply mathematical relationships or estimation to determine macroscopic variables for ideal gases.

**LEARNING OBJECTIVE 2.7** The student is able to explain how solutes can be separated by chromatography based on intermolecular interactions.

**LEARNING OBJECTIVE 2.8** The student can draw and/or interpret representations of solutions that show the interactions between the solute and solvent.

**LEARNING OBJECTIVE 2.9** The student is able to create or interpret representations that link the concept of molarity with particle views of solutions.

**LEARNING OBJECTIVE 2.10** The student can design and/or interpret the results of a separation experiment (filtration, paper chromatography, column chromatography, or distillation) in terms of the relative strength of interactions among and between the components.

**LEARNING OBJECTIVE 2.11** The student is able to explain the trends in properties and/or predict properties of samples consisting of particles with no permanent dipole on the basis of London dispersion forces.

**LEARNING OBJECTIVE 2.12** The student can qualitatively analyze data regarding real gases to identify deviations from ideal behavior and relate these to molecular interactions.

**LEARNING OBJECTIVE 2.13** The student is able to describe the relationships between the structural features of polar molecules and the forces of attraction between the particles.

**LEARNING OBJECTIVE 2.14** The student is able to apply Coulomb's law qualitatively (including using representations) to describe the interactions of ions, and the attractions between ions and solvents to explain the factors that contribute to the solubility of ionic compounds.

**LEARNING OBJECTIVE 2.15** The student is able to explain observations regarding the solubility of ionic solids and molecules in water and other solvents on the basis of particle views that include intermolecular interactions and entropic effects.

**LEARNING OBJECTIVE 2.16** The student is able to explain the properties (phase, vapor pressure, viscosity, etc.) of small and large molecular compounds in terms of the strengths and types of intermolecular forces.

**LEARNING OBJECTIVE 2.17** The student can predict the type of bonding present between two atoms in a binary compound based on position in the periodic table and the electronegativity of the elements.

**LEARNING OBJECTIVE 2.18** The student is able to rank and justify the ranking of bond polarity on the basis of the locations of the bonded atoms in the periodic table.

**LEARNING OBJECTIVE 2.19** The student can create visual representations of ionic substances that connect the microscopic structure to macroscopic properties, and/or use representations to connect the microscopic structure to macroscopic properties (e.g., boiling point,

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solubility, hardness, brittleness, low volatility, lack of malleability, ductility, or conductivity).

**LEARNING OBJECTIVE 2.20** The student is able to explain how a bonding model involving delocalized electrons is consistent with macroscopic properties of metals (e.g., conductivity, malleability, ductility, and low volatility) and the shell model of the atom.

**LEARNING OBJECTIVE 2.21** The student is able to use Lewis diagrams and VSEPR to predict the geometry of molecules, identify hybridization, and make predictions about polarity.

**LEARNING OBJECTIVE 2.22** The student is able to design or evaluate a plan to collect and/or interpret data needed to deduce the type of bonding in a sample of a solid.

**LEARNING OBJECTIVE 2.23** The student can create a representation of an ionic solid that shows essential characteristics of the structure and interactions present in the substance.

**LEARNING OBJECTIVE 2.24** The student is able to explain a representation that connects properties of an ionic solid to its structural attributes and to the interactions present at the atomic level.

**LEARNING OBJECTIVE 2.25** The student is able to compare the properties of metal alloys with their constituent elements to determine if an alloy has formed, identify the type of alloy formed, and explain the differences in properties using particulate level reasoning.

**LEARNING OBJECTIVE 2.26** Students can use the electron sea model of metallic bonding to predict or make claims about the macroscopic properties of metals or alloys.

**LEARNING OBJECTIVE 2.27** The student can create a representation of a metallic solid that shows essential characteristics of the structure and interactions present in the substance.

**LEARNING OBJECTIVE 2.28** The student is able to explain a representation that connects properties of a metallic solid to its structural attributes and to the interactions present at the atomic level.

**LEARNING OBJECTIVE 2.29** The student can create a representation of a covalent solid that shows essential characteristics of the structure and interactions present in the substance.

**LEARNING OBJECTIVE 2.30** The student is able to explain a representation that connects properties of a covalent solid to its structural attributes and to the interactions present at the atomic level.

**LEARNING OBJECTIVE 2.31** The student can create a representation of a molecular solid that shows essential characteristics of the structure and interactions present in the substance.

**LEARNING OBJECTIVE 2.32** The student is able to explain a representation that connects properties of a molecular solid to its structural attributes and to the interactions present at the atomic level.

**LEARNING OBJECTIVE 3.1** Students can translate among macroscopic observations of change, chemical equations, and particle views.

**LEARNING OBJECTIVE 3.2** The student can translate an observed chemical change into a balanced chemical equation and justify the choice of equation type (molecular, ionic, or net ionic) in terms of utility for the given circumstances.

**LEARNING OBJECTIVE 3.3** The student is able to use stoichiometric calculations to predict the results of performing a reaction in the laboratory and/or to analyze deviations from the expected results.

**LEARNING OBJECTIVE 3.4** The student is able to relate quantities (measured mass of substances, volumes of solutions, or volumes and pressures of gases) to identify stoichiometric relationships for a reaction, including situations involving limiting reactants and situations in which the reaction has not gone to completion.

**LEARNING OBJECTIVE 3.5** The student is able to design a plan in order to collect data on the synthesis or decomposition of a compound to confirm the conservation of matter and the law of definite proportions.

**LEARNING OBJECTIVE 3.6** The student is able to use data from synthesis or decomposition of a compound to confirm the conservation of matter and the law of definite proportions.

**LEARNING OBJECTIVE 3.7** The student is able to identify compounds as Bronsted-Lowry acids, bases, and/or conjugate acid-base pairs, using proton-transfer reactions to justify the identification.

**LEARNING OBJECTIVE 3.8** The student is able to identify redox reactions and justify the identification in terms of electron transfer.

**LEARNING OBJECTIVE 3.9** The student is able to design and/or interpret the results of an experiment involving a redox titration.

**LEARNING OBJECTIVE 3.10** The student is able to evaluate the classification of a process as a physical change, chemical change, or ambiguous change based on both macroscopic observations and the distinction between rearrangement of covalent interactions and noncovalent interactions.

**LEARNING OBJECTIVE 3.11** The student is able to interpret observations regarding macroscopic energy changes associated with a reaction or process to generate a relevant symbolic and/or graphical representation of the energy changes.

**LEARNING OBJECTIVE 3.12** The student can make qualitative or quantitative predictions about galvanic or electrolytic reactions based on half-cell reactions and potentials and/or Faraday's laws.

**LEARNING OBJECTIVE 3.13** The student can analyze data regarding galvanic or electrolytic cells to identify properties of the underlying redox reactions.

**LEARNING OBJECTIVE 4.1** The student is able to design and/or interpret the results of an experiment regarding the factors (i.e., temperature, concentration, surface area) that may influence the rate of a reaction.

**LEARNING OBJECTIVE 4.2** The student is able to analyze concentration vs. time data to determine the rate law for a zeroth-, first-, or second-order reaction.

**LEARNING OBJECTIVE 4.3** The student is able to connect the half-life of a reaction to the rate constant of a first-order reaction and justify the use of this relation in terms of the reaction being a first-order reaction.

**LEARNING OBJECTIVE 4.4** The student is able to connect the rate law for an elementary reaction to the frequency and success of molecular collisions, including connecting the frequency and success to the order and rate constant, respectively.

**LEARNING OBJECTIVE 4.5** The student is able to explain the difference between collisions that convert reactants to products and those that do not in terms of energy distributions and molecular orientation.

**LEARNING OBJECTIVE 4.6** The student is able to use representations of the energy profile for an elementary reaction (from the reactants, through the transition state, to the products) to make qualitative predictions regarding the relative temperature dependence of the reaction rate.

**LEARNING OBJECTIVE 4.7** The student is able to evaluate alternative explanations, as expressed by reaction mechanisms, to determine which are consistent with data regarding the overall rate of a reaction, and data that can be used to infer the presence of a reaction intermediate.

**LEARNING OBJECTIVE 4.8** The student can translate among reaction energy profile representations, particulate representations, and symbolic representations (chemical equations) of a chemical reaction occurring in the presence and absence of a catalyst.

**LEARNING OBJECTIVE 4.9** The student is able to explain changes in reaction rates arising from the use of acid-base catalysts, surface catalysts, or enzyme catalysts, including selecting appropriate mechanisms with or without the catalyst present.

**LEARNING OBJECTIVE 5.1** The student is able to create or use graphical representations in order to connect the dependence of potential energy to the distance between atoms and factors, such as bond order (for covalent interactions) and polarity (for intermolecular interactions), which influence the interaction strength.

**LEARNING OBJECTIVE 5.2** The student is able to relate temperature to the motions of particles, either via particulate representations, such as drawings of particles with arrows indicating velocities, and/or via representations of average kinetic energy and distribution of kinetic energies of the particles, such as plots of the Maxwell-Boltzmann distribution.

**LEARNING OBJECTIVE 5.3** The student can generate explanations or make predictions about the transfer of thermal energy between systems based on this transfer being due to a kinetic energy transfer between systems arising from molecular collisions.

**LEARNING OBJECTIVE 5.4** The student is able to use conservation of energy to relate the magnitudes of the energy changes occurring in two or more interacting systems, including identification of the systems, the type (heat versus work), or the direction of energy flow.



**LEARNING OBJECTIVE 5.5** The student is able to use conservation of energy to relate the magnitudes of the energy changes when two nonreacting substances are mixed or brought into contact with one another.

**LEARNING OBJECTIVE 5.6** The student is able to use calculations or estimations to relate energy changes associated with heating/cooling a substance to the heat capacity, relate energy changes associated with a phase transition to the enthalpy of fusion/vaporization, relate energy changes associated with a chemical reaction to the enthalpy of the reaction, and relate energy changes to  $P\Delta V$  work.

**LEARNING OBJECTIVE 5.7** The student is able to design and/or interpret the results of an experiment in which calorimetry is used to determine the change in enthalpy of a chemical process (heating/cooling, phase transition, or chemical reaction) at constant pressure.

**LEARNING OBJECTIVE 5.8** The student is able to draw qualitative and quantitative connections between the reaction enthalpy and the energies involved in the breaking and formation of chemical bonds.

**LEARNING OBJECTIVE 5.9** The student is able to make claims and/or predictions regarding relative magnitudes of the forces acting within collections of interacting molecules based on the distribution of electrons within the molecules and the types of intermolecular forces through which the molecules interact.

**LEARNING OBJECTIVE 5.10** The student can support the claim about whether a process is a chemical or physical change (or may be classified as both) based on whether the process involves changes in intramolecular versus intermolecular interactions.

**LEARNING OBJECTIVE 5.11** The student is able to identify the noncovalent interactions within and between large molecules, and/or connect the shape and function of the large molecule to the presence and magnitude of these interactions.

**LEARNING OBJECTIVE 5.12** The student is able to use representations and models to predict the sign and relative magnitude of the entropy change associated with chemical or physical processes.

**LEARNING OBJECTIVE 5.13** The student is able to predict whether or not a physical or chemical process is thermodynamically favored by determination of (either quantitatively or qualitatively) the signs of both  $\Delta H^\circ$  and  $\Delta S^\circ$ , and calculation or estimation of  $\Delta G^\circ$  when needed.

**LEARNING OBJECTIVE 5.14** The student is able to determine whether a chemical or physical process is thermodynamically favorable by calculating the change in standard Gibbs free energy.

**LEARNING OBJECTIVE 5.15** The student is able to explain how the application of external energy sources or the coupling of favorable with unfavorable reactions can be used to cause processes that are not thermodynamically favorable to become favorable.

**LEARNING OBJECTIVE 5.16** The student can use LeChâtelier's principle to make qualitative predictions for systems in which coupled reactions that share a common intermediate drive formation of a product.

**LEARNING OBJECTIVE 5.17** The student can make quantitative predictions for systems involving coupled reactions that share a common intermediate, based on the equilibrium constant for the combined reaction.

**LEARNING OBJECTIVE 5.18** The student can explain why a thermodynamically favored chemical reaction may not produce large amounts of product (based on consideration of both initial conditions and kinetic effects), or why a thermodynamically unfavored chemical reaction can produce large amounts of product for certain sets of initial conditions.

**LEARNING OBJECTIVE 6.1** The student is able to, given a set of experimental observations regarding physical, chemical, biological, or environmental processes that are reversible, construct an explanation that connects the observations to the reversibility of the underlying chemical reactions or processes.

**LEARNING OBJECTIVE 6.2** The student can, given a manipulation of a chemical reaction or set of reactions (e.g., reversal of reaction or addition of two reactions), determine the effects of that manipulation on  $Q$  or  $K$ .

**LEARNING OBJECTIVE 6.3** The student can connect kinetics to equilibrium by using reasoning about equilibrium, such as LeChâtelier's principle, to infer the relative rates of the forward and reverse reactions.

**LEARNING OBJECTIVE 6.4** The student can, given a set of initial conditions (concentrations or partial pressures) and the equilibrium constant,  $K$ , use the tendency of  $Q$  to approach  $K$  to predict and justify the prediction as to whether the reaction will proceed toward products or reactants as equilibrium is approached.

**LEARNING OBJECTIVE 6.5** The student can, given data (tabular, graphical, etc.) from which the state of a system at equilibrium can be obtained, calculate the equilibrium constant,  $K$ .

**LEARNING OBJECTIVE 6.6** The student can, given a set of initial conditions (concentrations or partial pressures) and the equilibrium constant,  $K$ , use stoichiometric relationships and the law of mass action ( $Q$  equals  $K$  at equilibrium) to determine qualitatively and/or quantitatively the conditions at equilibrium for a system involving a single reversible reaction.

**LEARNING OBJECTIVE 6.7** The student is able, for a reversible reaction that has a large or small  $K$ , to determine which chemical species will have very large versus very small concentrations at equilibrium.

**LEARNING OBJECTIVE 6.8** The student is able to use LeChâtelier's principle to predict the direction of the shift resulting from various possible stresses on a system at chemical equilibrium.

**LEARNING OBJECTIVE 6.9** The student is able to use LeChâtelier's principle to design a set of conditions that will optimize a desired outcome, such as product yield.

**LEARNING OBJECTIVE 6.10** The student is able to connect LeChâtelier's principle to the comparison of  $Q$  to  $K$  by explaining the effects of the stress on  $Q$  and  $K$ .

**LEARNING OBJECTIVE 6.11** The student can generate or use a particulate representation of an acid (strong or weak or polyprotic) and a strong base to explain the species that will have large versus small concentrations at equilibrium.

**LEARNING OBJECTIVE 6.12** The student can reason about the distinction between strong and weak acid solutions with similar values of pH, including the percent ionization of the acids, the concentrations needed to achieve the same pH, and the amount of base needed to reach the equivalence point in a titration.

**LEARNING OBJECTIVE 6.13** The student can interpret titration data for monoprotic or polyprotic acids involving titration of a weak or strong acid by a strong base (or a weak or strong base by a strong acid) to determine the concentration of the titrant and the  $pK_a$  for a weak acid, or the  $pK_b$  for a weak base.

**LEARNING OBJECTIVE 6.14** The student can, based on the dependence of  $K_w$  on temperature, reason that neutrality requires  $[H^+] = [OH^-]$  as opposed to requiring  $pH = 7$ , including especially the applications to biological systems.

**LEARNING OBJECTIVE 6.15** The student can identify a given solution as containing a mixture of strong acids and/or bases and calculate or estimate the pH (and concentrations of all chemical species) in the resulting solution.

**LEARNING OBJECTIVE 6.16** The student can identify a given solution as being the solution of a monoprotic weak acid or base (including salts in which one ion is a weak acid or base), calculate the pH and concentration of all species in the solution, and/or infer the relative strengths of the weak acids or bases from given equilibrium concentrations.

**LEARNING OBJECTIVE 6.17** The student can, given an arbitrary mixture of weak and strong acids and bases (including polyprotic systems), determine which species will react strongly with one another (i.e., with  $K > 1$ ) and what species will be present in large concentrations at equilibrium.

**LEARNING OBJECTIVE 6.18** The student can design a buffer solution with a target pH and buffer capacity by selecting an appropriate conjugate acid-base pair and estimating the concentrations needed to achieve the desired capacity.

**LEARNING OBJECTIVE 6.19** The student can relate the predominant form of a chemical species involving a labile proton (i.e., protonated/deprotonated form of a weak acid) to the pH of a solution and the  $pK_a$  associated with the labile proton.

**LEARNING OBJECTIVE 6.20** The student can identify a solution as being a buffer solution and explain the buffer mechanism in terms of the reactions that would occur on addition of acid or base.

**LEARNING OBJECTIVE 6.21** The student can predict the solubility of a salt, or rank the solubility of salts, given the relevant  $K_{sp}$  values.

**LEARNING OBJECTIVE 6.22** The student can interpret data regarding solubility of salts to determine, or rank, the relevant  $K_{sp}$  values.

**LEARNING OBJECTIVE 6.23** The student can interpret data regarding the relative solubility of salts in terms of factors (common ions, pH) that influence the solubility.

**LEARNING OBJECTIVE 6.24** The student can analyze the enthalpic and entropic changes associated with the dissolution of a salt, using particulate level interactions and representations.

**LEARNING OBJECTIVE 6.25** The student is able to express the equilibrium constant in terms of  $\Delta G^\circ$  and  $RT$  and use this relationship to estimate the magnitude of  $K$  and, consequently, the thermodynamic favorability of the process.

# A Diagnostic Test

The purpose of this test is to give you an indication of how well you will perform on the AP Chemistry exam. These questions are representative of the AP Chemistry examination, but bear in mind it is impossible to predict exactly how well you will do on the actual exam. Calculators may not be used for answering questions in the first section of this test. The first section is 50% of your total test grade. Time yourself to finish this part in 90 minutes. There are two types of multiple-choice questions used in this examination. One type consists of a data set that will be used to answer 3–5 multiple-choice questions. The other type of multiple-choice question consists of a question or incomplete statements followed by four possible answers. Select the one that is best in each case.

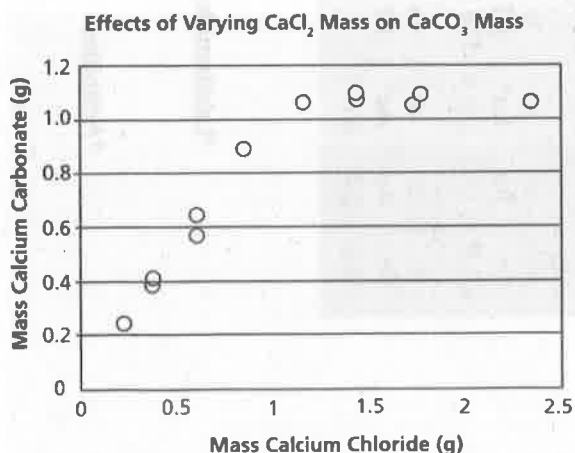
## AP CHEMISTRY EXAMINATION Section I: Multiple-Choice Questions Time: 90 minutes Number of Questions: 60

No calculators can be used in this section. A periodic table and a formula chart with constants is provided.

**Directions:** Each of the questions or incomplete statements below is followed by four suggested answers or completions. Select the one that is best in each case.

Questions 1–3 refer to the following:

Various masses of calcium chloride were dissolved in water and added to 6.00 mL of 1.7 M sodium carbonate. The graph of the data is below.



1. What technique would separate calcium carbonate precipitate from the rest of the substances in the reaction?  
(A) column chromatography  
(B) distillation  
(C) evaporation  
(D) filtration
2. Which chemical was NOT present in the reaction vessel after 0.5 grams of calcium chloride reacted?  
(A) calcium chloride  
(B) sodium carbonate  
(C) calcium carbonate  
(D) sodium chloride

**Periodic Table of Elements**

1 H 1.008																	2 He 4.003							
3 Li 6.941	4 Be 9.012															9 F 19.00	10 Ne 20.18							
11 Na 22.99	12 Mg 24.31															17 Cl 35.45	18 Ar 39.95							
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80							
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3							
55 Cs 132.9	56 Ba 137.3	57 La* 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)							
87 Fr (223)	88 Ra 226	89 Ac† (227)	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg (272)														
											65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0							
											90 Th 232.0	91 Pa (231)	92 U 238.0	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

\*Lanthanides

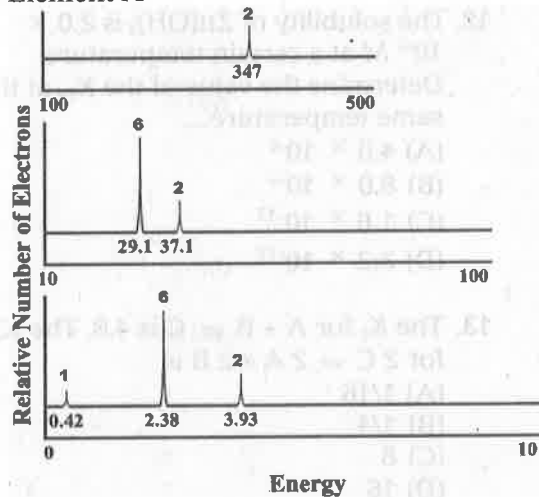
†Actinides

3. Explain the relatively horizontal portion of the graph from about 1.2 g to 2.4 g of  $\text{CaCl}_2$ .
- (A) Calcium chloride was the limiting reagent.
  - (B) Sodium carbonate was the limiting reagent.
  - (C) Calcium carbonate was no longer being produced.
  - (D) Calcium chloride and sodium carbonate reached equilibrium.

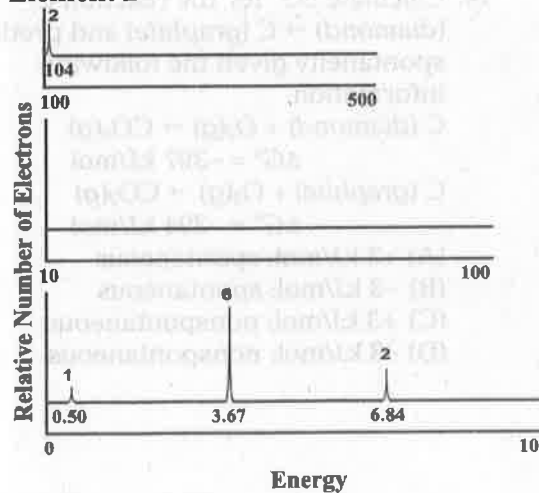
Questions 4–7 refer to the following:

Photoelectron spectra of 2 elements are shown below. The units of energy are MJ.

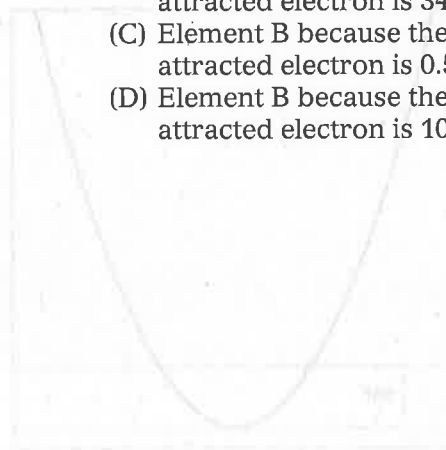
Element A



Element B



4. How many shells or energy levels are occupied in Element A?
- (A) 3
  - (B) 4
  - (C) 5
  - (D) 6
5. Which electron is in the valence shell of Element A?
- (A) 0.42
  - (B) 3.93
  - (C) 37.1
  - (D) 347
6. Propose a reason why the peak of highest energy for Element A has a higher energy (347 MJ) than the peak of highest energy in Element B (104 MJ).
- (A) The electrons in Element A are farther from the nucleus than the electrons in Element B.
  - (B) The electrons in Element A are closer to the nucleus than the electrons in Element B.
  - (C) Element A has more electrons than Element B.
  - (D) Element A has greater shielding than Element B.
7. Which element would be more reactive with water?
- (A) Element A because the least attracted electron is 0.42 MJ.
  - (B) Element A because the most attracted electron is 347 MJ.
  - (C) Element B because the least attracted electron is 0.50 MJ.
  - (D) Element B because the most attracted electron is 104 MJ.



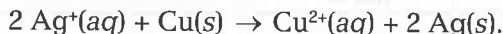
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### 32 ❖ A DIAGNOSTIC TEST

For questions 8–9, consider the electrochemical cell:



The cell reaction is



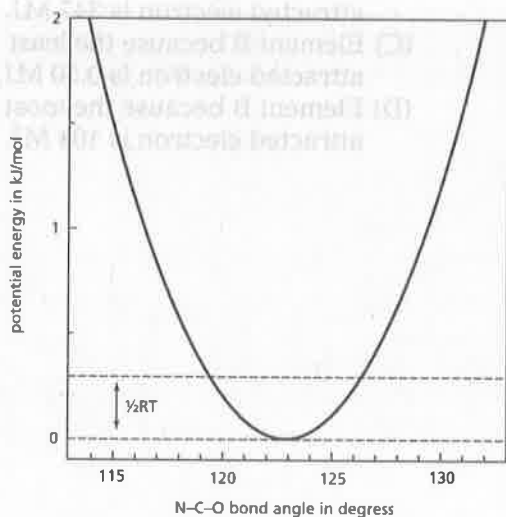
The measured voltage is +0.46 V.

8. Increasing the concentration of silver ions will
- cause a decrease in blue color in the cell
  - increase the cell voltage above +0.46 V
  - decrease the concentration of copper(II) ions
  - cause no change in the cell voltage

9. The reaction at the anode is

- $\text{Ag}^{+}(\text{aq}) + \text{e}^{-} \rightarrow \text{Ag}(\text{s})$
- $\text{Ag}(\text{s}) \rightarrow \text{Ag}^{+}(\text{aq})$
- $\text{Cu}(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-}$
- $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Cu}(\text{s})$

10. The diagram below shows how the potential energy of N–C–O bond is related to bond angle. At 300 K, the atoms vibrate and deviate approximately  $\pm 4^{\circ}$  from the average bond angle. How does an increase in temperature affect the deviation in the bond angle?



- No change because temperature does not affect bond angle.
- Increase because molecules are moving faster.
- Increase because atoms have more kinetic energy.
- Increase because as the degrees of temperature increase, the degrees of angle increase.

11. The amount of silver which will be formed when 0.00200 mol of  $\text{Ag}_2\text{S}$  reacts completely with excess zinc is
- 0.00100 mol
  - 0.00200 mol
  - 0.00400 mol
  - 0.00800 mol

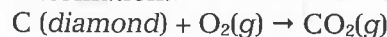
12. The solubility of  $\text{Zn}(\text{OH})_2$  is  $2.0 \times 10^{-6}\text{ M}$  at a certain temperature. Determine the value of the  $K_{\text{sp}}$  at this same temperature.

- $4.0 \times 10^{-6}$
- $8.0 \times 10^{-6}$
- $1.6 \times 10^{-17}$
- $3.2 \times 10^{-17}$

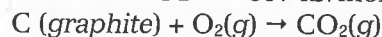
13. The  $K_c$  for  $\text{A} + \text{B} \rightleftharpoons \text{C}$  is 4.0. The  $K_c$  for  $2\text{C} \rightleftharpoons 2\text{A} + 2\text{B}$  is

- 1/16
- 1/4
- 8
- 16

14. Calculate  $\Delta G^{\circ}$  for the reaction  $\text{C}(\text{diamond}) \rightarrow \text{C}(\text{graphite})$  and predict spontaneity given the following information.



$$\Delta G^{\circ} = -397\text{ kJ/mol}$$



$$\Delta G^{\circ} = -394\text{ kJ/mol}$$

- +3 kJ/mol; spontaneous
- 3 kJ/mol; spontaneous
- +3 kJ/mol; nonspontaneous
- 3 kJ/mol; nonspontaneous

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Calcium Oxide



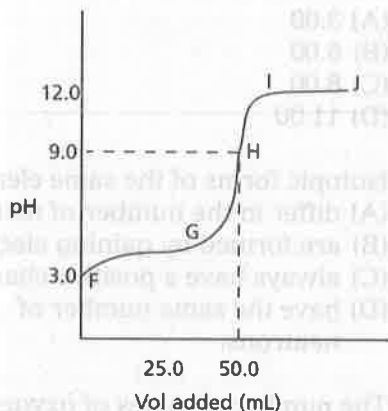
Magnesium Oxide

15. Calcium oxide,  $\text{CaO}$ , has a lower melting temperature than magnesium oxide,  $\text{MgO}$ , due to the
- higher charge density of  $\text{Mg}^{2+}$  than of  $\text{Ca}^{2+}$
  - higher charge density of  $\text{Ca}^{2+}$  than of  $\text{O}^{2-}$
  - greater atomic volume of  $\text{Mg}^{2+}$  than of  $\text{O}^{2-}$
  - greater atomic volume of  $\text{Mg}^{2+}$  than of  $\text{Ca}^{2+}$
16. The equilibrium system not affected by the pressure change which results from a volume change at constant temperature is
- $2 \text{O}_3(g) \rightleftharpoons 3 \text{O}_2(g)$
  - $\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g)$
  - $\text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2 \text{HI}(g)$
  - $2 \text{NO}_2(g) \rightleftharpoons \text{N}_2\text{O}_4(g)$
17. A solution of 50.0 mL of 0.0010 M  $\text{Ba}(\text{OH})_2$  is slowly titrated with 50.0 mL of 0.0030 M  $\text{H}_2\text{SO}_4$ . The conductivity of this solution will
- decrease to near zero, then increase
  - decrease to near zero and remain very low
  - increase as the acid is added then become constant at a high value
  - increase as the acid is added and then slowly become very low
18. Which of the following pairs illustrates the law of multiple proportions?
- $\text{SO}_2$ ,  $\text{SO}_3$
  - $\text{CO}_2$ ,  $\text{CCl}_4$
  - $\text{NaCl}$ ,  $\text{NaBr}$
  - $\text{NH}_4\text{Cl}$ ,  $\text{NH}_4\text{Br}$
19. Determine the pH of a 0.10 M solution of a base, B, with a  $K_b$  of  $1.0 \times 10^{-5}$ .
- 3.00
  - 6.00
  - 8.00
  - 11.00
20. Isotopic forms of the same element
- differ in the number of neutrons
  - are formed by gaining electrons
  - always have a positive charge
  - have the same number of neutrons
21. The number of moles of oxygen atoms in one mole of iron(II) phosphate is
- 2
  - 3
  - 4
  - 8
22. Magnesium fluoride, a salt of low solubility in water, has a  $K_{sp}$  of  $6.4 \times 10^{-9}$ . The concentration of  $\text{Mg}^{2+}$  ions in this solution would be
- $\sqrt{6.4 \times 10^{-9}} \text{ M}$
  - $\sqrt{(6.4 \times 10^{-9} / 2)} \text{ M}$
  - $\sqrt[3]{(6.4 \times 10^{-9} / 3)} \text{ M}$
  - $\sqrt[3]{(6.4 \times 10^{-9} / 4)} \text{ M}$

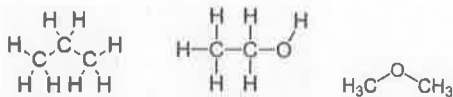
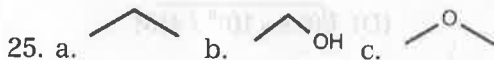
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34 ❖ A DIAGNOSTIC TEST

Questions 23 and 24 refer to this pH curve for the titration of 50.0 mL of 0.100 M acid with 0.100 M base.



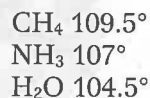
23. What type of titration does this curve represent?
- strong acid/strong base
  - strong acid/weak base
  - weak acid/strong base
  - either a strong acid/strong base or a weak acid/strong base
24. Buffering is most effective
- between points F and G
  - between points G and H
  - at point H between points I and J
  - between both F and G and between I and J



Arrange the above liquids in order of increasing vapor pressure:

- $b < c < a$
- $b < a < c$
- $a < c < b$
- $a < b < c$

26. Bond angle data shows the following angles for three compounds:

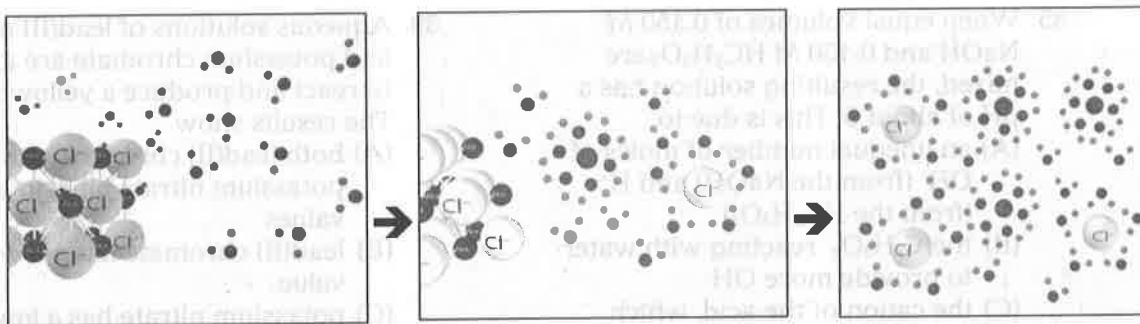


This trend is basically because

- lone pairs of electrons require more room than bonding pairs
  - hydrogen atoms repel each other more in water than in  $\text{CH}_4$  or  $\text{NH}_3$
  - oxygen has a higher electronegativity than does N, and C has even less
  - of the attempt of all central atoms to achieve the tetrahedral shape
27. Which of the following are in order of increasing boiling points?
- $\text{RbCl} < \text{CH}_3\text{Cl} < \text{CH}_3\text{OH} < \text{CH}_4$
  - $\text{CH}_4 < \text{CH}_3\text{Cl} < \text{CH}_3\text{OH} < \text{RbCl}$
  - $\text{CH}_4 < \text{CH}_3\text{OH} < \text{CH}_3\text{Cl} < \text{RbCl}$
  - $\text{RbCl} < \text{CH}_3\text{OH} < \text{CH}_3\text{Cl} < \text{CH}_4$

28. When 100.0 mL of 2.0 M  $\text{NH}_3$  and 100.0 mL of 1.0 M  $\text{AgNO}_3$  are mixed, but before any reaction occurs, the major species in solution are
- $\text{Ag}^+$ ,  $\text{NO}_3^-$ ,  $\text{NH}_3$ , and  $\text{H}_2\text{O}$
  - $\text{Ag}^+$ ,  $\text{NO}_3^-$ ,  $\text{NH}_4^+$ , and  $\text{OH}^-$
  - $\text{NH}_3$  and  $\text{Ag}(\text{NH}_3)_2^+$
  - $\text{NH}_3$ ,  $\text{Ag}(\text{NH}_3)_2^+$ , and  $\text{H}_2\text{O}$

29. The geometry of a  $\text{PH}_3$  molecule is described by the VSEPR model as
- trigonal planar
  - tetrahedral
  - bent or angular
  - trigonal pyramidal



30. The dissolution of sodium chloride shown in the diagram above may be considered either a chemical or a physical change. Justify this statement.

- |     | <b>Chemical Change</b>                                                    | <b>Physical Change</b>                                                    |
|-----|---------------------------------------------------------------------------|---------------------------------------------------------------------------|
| (A) | Na <sup>+</sup> and Cl <sup>-</sup> bonds are broken.                     | H-O-H bonds are broken.                                                   |
| (B) | Na <sup>+</sup> and Cl <sup>-</sup> form ion-dipole attractions to water. | Hydrogen bonds between water molecules are broken.                        |
| (C) | H-O-H bonds are broken.                                                   | Na <sup>+</sup> and Cl <sup>-</sup> bonds are broken.                     |
| (D) | Hydrogen bonds between water molecules are broken.                        | Na <sup>+</sup> and Cl <sup>-</sup> form ion-dipole attractions to water. |

31. The molecule whose Lewis structure requires resonance structures to best explain its bonding is

- (A) CO<sub>2</sub>  
 (B) PCl<sub>5</sub>  
 (C) OF<sub>5</sub>  
 (D) SO<sub>2</sub>

32.  $A + B \rightleftharpoons 2C$

The equilibrium concentrations of reactants and products for the reaction above are

Substance	Concentration (M)
A	0.25
B	0.25
C	0.050

What is the value of the equilibrium constant,  $K$ ?

- (A) 0.040  
 (B) 0.16  
 (C) 0.80  
 (D) 25

33. The  $K_a$  for a weak acid is  $5.0 \times 10^{-10}$  at 25°C. Determine the value of  $K_b$  for the conjugate base of this weak acid.

- (A)  $0.50 \times 10^{-5}$   
 (B)  $1.5 \times 10^{-5}$   
 (C)  $2.0 \times 10^{-5}$   
 (D)  $5.0 \times 10^{-5}$

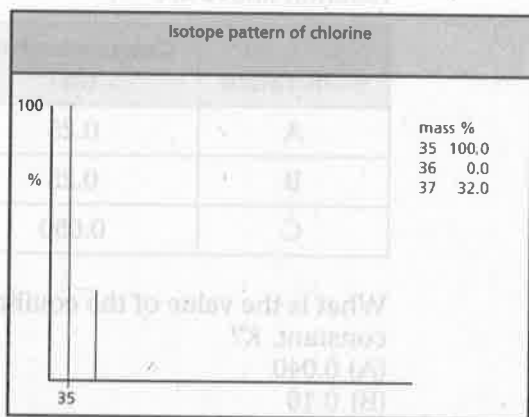
34. The species with the most polar bond is

- (A) F-F  
 (B) Sc-Ti  
 (C) Cr-Br  
 (D) P-Cl

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36 ❖ A DIAGNOSTIC TEST

35. When equal volumes of 0.150 M NaOH and 0.150 M HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub> are mixed, the resulting solution has a pH of about 9. This is due to
- an unequal number of moles of OH<sup>-</sup> (from the NaOH) and H<sup>+</sup> (from the HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub>)
  - the C<sub>2</sub>H<sub>3</sub>O<sub>2</sub><sup>-</sup> reacting with water to provide more OH<sup>-</sup>
  - the cation of the acid, which remains in solution at the equivalent point, and is a base
  - both the C<sub>2</sub>H<sub>3</sub>O<sub>2</sub><sup>-</sup> reacting with water and the acidic nature of the anion of the acid

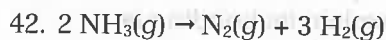


36. According to the mass spectrum above, the isotopic composition of <sup>35</sup>Cl:<sup>37</sup>Cl is closest to
- 35:37
  - 100:32
  - 50:50
  - 75:25
37. Of the following, the species that has the largest radius is
- Cl<sup>-</sup>
  - Br<sup>-</sup>
  - K<sup>+</sup>
  - Sr<sup>2+</sup>
38. The pH of a 0.001 M HBr solution is
- 3.0
  - 7.0
  - 11.0
  - impossible to determine without more data

39. Aqueous solutions of lead(II) nitrate and potassium chromate are allowed to react and produce a yellow solid. The results show
- both lead(II) chromate and potassium nitrate have low K<sub>sp</sub> values
  - lead(II) chromate has a low K<sub>sp</sub> value
  - potassium nitrate has a low K<sub>sp</sub> value
  - potassium chromate has a low K<sub>sp</sub> value

40. A rigid cylinder contains CO<sub>2</sub> gas at a constant temperature. Some of the carbon dioxide is allowed to escape. Which of the following applies to the CO<sub>2</sub>?
- The pressure of the gas increases.
  - The volume of the gas decreases.
  - The average molecular speed decreases.
  - The distance between CO<sub>2</sub> molecules is increased.

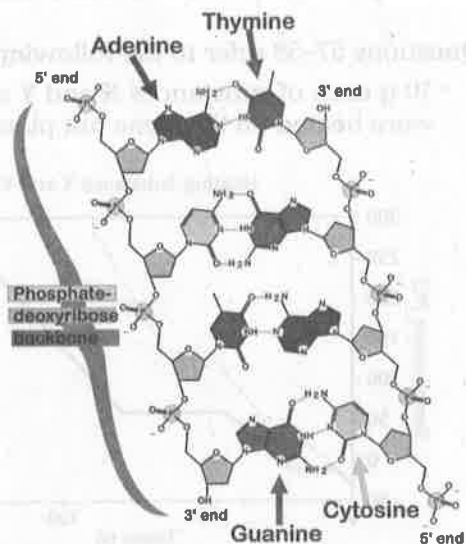
41. Which of the following is most likely to be a brittle compound with low conductivity as a solid, and have a high melting point?
- RbF
  - CCl<sub>4</sub>
  - ICl
  - SF<sub>6</sub>



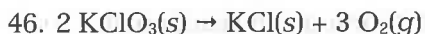
The above reaction occurs in a closed system of constant volume and temperature. What is the resultant pressure of the hydrogen if the partial pressure of the ammonia decreases by 0.40 atm?

- increases by 0.20 atm
- increases by 0.40 atm
- increases by 0.60 atm
- decreases by 0.60 atm

43. Air is pumped into a rigid steel cylinder at constant temperature. The increase in pressure is due to  
 (A) increased molecular collisions  
 (B) the greater kinetic energy of the gas particles  
 (C) increase in the size of the individual molecules  
 (D) the greater force of attraction between gas molecules at high pressure
44. Which of the following gases would depart the most from ideal behavior?  
 (A)  $H_2$   
 (B) Xe  
 (C) He  
 (D)  $N_2$



45. The double helix structure of deoxyribonucleic acid (DNA) is shown above. When DNA replicates, the two strands must “unzip.” Which statement describes the energy requirements and forces involved?  
 (A) Energy must be added to break the covalent bonds.  
 (B) Energy must be added to break the hydrogen bonds.  
 (C) Energy is released when the covalent bonds break.  
 (D) Energy is released when the hydrogen bonds break.



According to the above equation, 0.40 mol of solid  $\text{KClO}_3$  completely decomposes, forming  $\text{KCl}$  and  $\text{O}_2$ . The dry gas is collected at STP. The volume of this oxygen gas would be most nearly  
 (A) 1.4 L  
 (B) 14 L  
 (C) 140 L  
 (D) 1400 L

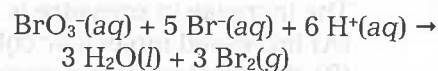
47. A hollow steel cylinder of volume 24 L contains 1.0 mole of  $N_2$  and 2.0 mol of Ar. The partial pressure of the  $N_2$  is  
 (A) 1/2 the total pressure  
 (B) 1/3 the total pressure  
 (C) 2 times that of the Ar  
 (D) equal to that of the Ar
48. As the atomic number increases from lithium to fluorine, the atomic radii  
 (A) do not change because the electrons are being added to the same energy level  
 (B) increase due to greater electron repulsion  
 (C) decrease because the increased the greater nuclear charge exerts a larger attraction on the electrons  
 (D) decrease because two orbitals (s and p) are filling

49. Assume that  $2A + B \rightarrow C$  is the rate determining step. 3.0 moles of A and 2.0 moles of B are placed in a 1.0-L flask; after five minutes the concentration of C reaches 1.0 M and the rate will have  
 (A) decreased by a factor of 6  
 (B) decreased by a factor of 9  
 (C) decreased by a factor of 10  
 (D) decreased by a factor of 18

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50. For a given reaction at a temperature of  $27^{\circ}\text{C}$ , the rate law is  $\text{Rate} = k [\text{X}][\text{Y}]$ . If the concentration of X and of Y are both  $0.40 \text{ M}$ , the rate is  $4.0 \times 10^{-6} \text{ mol L}^{-1} \text{ min}^{-1}$ . Determine the value of  $k$  (the rate constant) at this temperature.
- (A)  $2.5 \times 10^{-5} \text{ L mol}^{-1} \text{ min}^{-1}$   
 (B)  $2.5 \times 10^{-3} \text{ L mol}^{-1} \text{ min}^{-1}$   
 (C)  $2.5 \times 10^{-5} \text{ mol L}^{-1} \text{ min}^{-1}$   
 (D)  $1.0 \times 10^{-6} \text{ L mol}^{-1} \text{ min}^{-1}$
51. The major reason an increase in temperature causes an increase in reaction rate is that
- (A) the activation energy changes with temperature  
 (B) the fraction of high energy molecules increases  
 (C) molecules collide with greater frequency  
 (D) catalysts become more effective
52. Catalysts effectively increase reaction rate by
- (A) increasing the  $K_{\text{eq}}$   
 (B) increasing the concentration of the reactant  
 (C) decreasing the concentration of the products  
 (D) lowering the activation energy requirements
53. For all zero-order reactions,
- (A) the reaction rate is independent of time  
 (B) the rate constant equals zero  
 (C) the concentration of reactants does not change over time  
 (D) activation energy is very low
54. The first-order rate constant for nuclear decay of  $^{60}\text{Co}$  is  $0.13 \text{ yr}^{-1}$ , and for  $^{90}\text{Sr}$  it is  $0.24 \text{ yr}^{-1}$ .
- (A) The half-life of Sr is longer than that of Co.  
 (B) The half-life of Sr is shorter than that of Co.  
 (C) The half-lives of Sr and Co are equal.  
 (D) The half-lives of Sr and Co cannot be compared from these data.

55. To determine the order with respect to  $\text{Br}^-$  in the reaction



solutions should be prepared which differ in

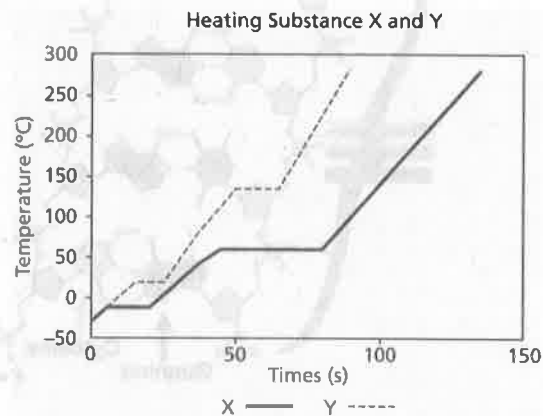
- (A)  $[\text{BrO}_3^-]$   
 (B)  $[\text{Br}^-]$   
 (C)  $[\text{H}^+]$   
 (D)  $[\text{BrO}_3^-]$  and  $[\text{Br}^-]$  and  $[\text{H}^+]$

56. Identify the process that produces the greatest change in entropy per mole of substance.

- (A)  $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$   
 (B)  $\text{CH}_3\text{OH}(\text{l}) \rightarrow \text{CH}_3\text{OH}(\text{aq})$   
 (C)  $\text{C}_{10}\text{H}_8(\text{s}) \rightarrow \text{C}_{10}\text{H}_8(\text{l})$   
 (D)  $\text{C}(\text{graphite}) \rightarrow \text{C}(\text{diamond})$

Questions 57–58 refer to the following.

10 g each of substances X and Y at  $-25^{\circ}\text{C}$  were heated on the same hot plate.



57. Identify and justify the substance with the higher specific heat capacity.
- (A) X; lower melting and boiling points  
 (B) X; more energy needed to melt and boil  
 (C) Y; higher melting and boiling points  
 (D) Y; less energy needed to melt and boil

58. After 50 s, which substance gained more heat?
- (A) Y because the temperature is higher
  - (B) X because it is boiling
  - (C) Y because its temperature is still changing
  - (D) They both absorbed the same amount of heat because the hot plate is the same
59. Which data must be collected to calculate the enthalpy of reaction?
- (A) volume, pressure, initial and final temperature
  - (B) conductivity, concentration, pH
  - (C) initial and final temperature, absorbance, concentration
  - (D) concentration, volume, initial and final temperature

Questions 59–60 refer to the following.

The enthalpy of reaction of HCl and NaOH is to be calculated. The solutions are combined in an insulated container that is isolated from the environment.

60. How does the temperature change and why does it occur?
- (A) The temperature increases because potential energy of the chemical bonds is transformed into thermal energy.
  - (B) The temperature increases due to the friction of mixing.
  - (C) The temperature decreases because water formed in the reaction has a high specific heat capacity.
  - (D) The temperature decreases due to decrease in entropy from the formation of a salt.

## Advanced Placement Chemistry Equations and Constants

Throughout the test the following symbols have the definitions specified unless otherwise noted.

L, mL = liter(s), milliliter(s)

g = gram(s)

nm = nanometer(s)

atm = atmosphere(s)

mm Hg = millimeters of mercury

J, kJ = joule(s), kilojoule(s)

V = volt(s)

mol = mole(s)

### ATOMIC STRUCTURE

$$E = h\nu$$

$$c = \lambda\nu$$

$E$  = energy

$\nu$  = frequency

$\lambda$  = wavelength

Planck's constant,  $h = 6.626 \times 10^{-34}$  J s

Speed of light,  $c = 2.998 \times 10^8$  m s<sup>-1</sup>

Avogadro's number =  $6.022 \times 10^{23}$  mol<sup>-1</sup>

Electron charge,  $e = -1.602 \times 10^{-19}$  coulomb

### EQUILIBRIUM

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}, \text{ where } a A + b B \rightleftharpoons c C + d D$$

$$K_p = \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b}$$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$K_b = \frac{[OH^-][HB^+]}{[B]}$$

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

$$= K_a \times K_b$$

$$\text{pH} = -\log[H^+], \text{ pOH} = -\log[OH^-]$$

$$14 = \text{pH} + \text{pOH}$$

$$\text{pH} = \text{p}K_a + \log \frac{[A^-]}{[HA]}$$

$$\text{p}K_a = -\log K_a, \text{ p}K_b = -\log K_b$$

#### Equilibrium Constants

$K_c$  (molar concentrations)

$K_p$  (gas pressures)

$K_a$  (weak acid)

$K_b$  (weak base)

$K_w$  (water)

### KINETICS

$$\ln[A]_t - \ln[A]_0 = -kt$$

$$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$$

$$t_{1/2} = \frac{0.693}{k}$$

$k$  = rate constant

$t$  = time

$t_{1/2}$  = half-life



**GASES, LIQUIDS, AND SOLUTIONS**

$$PV = nRT$$

$$P_A = P_{\text{total}} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{\text{total}} = P_A + P_B + P_C + \dots$$

$$n = \frac{m}{M}$$

$$K = ^\circ\text{C} + 273$$

$$D = \frac{m}{V}$$

$$KE \text{ per molecule} = \frac{1}{2}mv^2$$

Molarity,  $M$  = moles of solute per liter of solution

$$A = abc$$

$P$  = pressure

$V$  = volume

$T$  = temperature

$n$  = number of moles

$m$  = mass

$M$  = molar mass

$D$  = density

$KE$  = kinetic energy

$v$  = velocity

$A$  = absorbance

$a$  = molar absorptivity

$b$  = path length

$c$  = concentration

Gas constant,  $R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$

$$= 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1}$$

$$= 62.36 \text{ L torr mol}^{-1} \text{ K}^{-1}$$

$$1 \text{ atm} = 760 \text{ mm Hg}$$

$$= 760 \text{ torr}$$

STP =  $0.00^\circ\text{C}$  and  $1.000 \text{ atm}$

**THERMOCHEMISTRY/ ELECTROCHEMISTRY**

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ_{\text{products}} - \sum S^\circ_{\text{reactants}}$$

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

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$= -RT \ln K$$

$$= -nFE^\circ$$

$$I = \frac{q}{t}$$

$q$  = heat

$m$  = mass

$c$  = specific heat capacity

$T$  = temperature

$S^\circ$  = standard entropy

$H^\circ$  = standard enthalpy

$G^\circ$  = standard free energy

$n$  = number of moles

$E^\circ$  = standard reduction potential

$I$  = current (amperes)

$q$  = charge (coulombs)

$t$  = time (seconds)

Faraday's constant,  $F = 96,485$  coulombs per mole of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

**GO ON TO NEXT PAGE**

## Introduction to Section II: Free-Response Questions

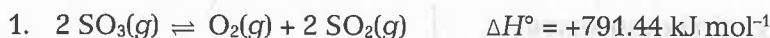
Section II of the AP Chemistry Examination counts for 50% of the total test grade and involves several parts. Answering these questions gives you an opportunity to demonstrate your ability to present your material in clear, orderly, and convincing language. Your answers will be graded on the basis of accuracy, the kinds of information you include to support your responses, and the importance of the descriptive material used. Be specific; general, all-encompassing answers will not be graded as well as detailed answers with examples and equations. **CLEARLY SHOW THE METHOD USED AND THE STEPS INVOLVED IN ARRIVING AT YOUR ANSWERS.** It is to your advantage to do this, since you may obtain partial credit if you do and you will receive little or no credit if you do not. Attention should be paid to significant figures. On the AP exam, be sure to write all your answers to the questions on the lined pages following each question in the test booklet. Do not write your answers in the white space between questions.

### Section II: Free-Response Questions

**Time: 90 minutes**

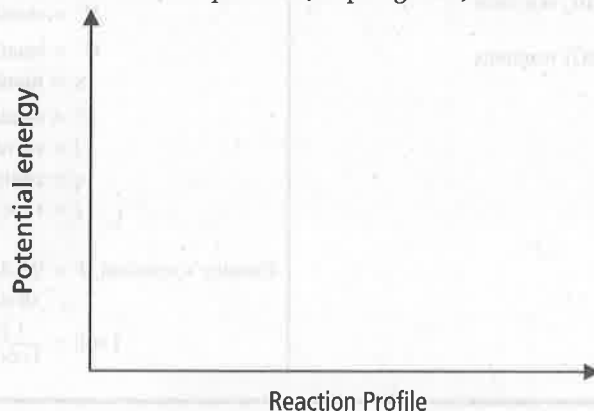
**Number of Questions: 7**

Allow yourself no more than 90 minutes to answer these questions. You may use a calculator, the equations sheet, and the periodic table throughout this section. All questions must be answered.



A 3.21-g sample of sulfur trioxide is placed in a 2.25-L cylinder and allowed to reach equilibrium at a constant temperature of 500. K, as shown in the above equation. Analysis shows  $1.23 \times 10^{-2}$  mol of sulfur dioxide at equilibrium.

(a) Sketch the reaction profile (or progress) on the axes given.



- (b) Write the equilibrium constant expression ( $K_c$ ) for this system.  
 (c) Calculate the concentration of all three gases at equilibrium.  
 (d) Calculate the  $K_c$  value for this system.  
 (e) Draw a particulate model with 10 particles to show the chemical species present in the cylinder at equilibrium. Clearly identify each chemical species in the model.

- (f) In a separate experiment, all of the variables remained the same except the temperature was raised to 750. K
- Predict how the equilibrium concentrations of the reactants and products will change.
  - Use bonding and collision theories to explain why the concentrations changed.

2. Perbromic acid,  $\text{HBrO}_4$ , can react with sulfuric acid,  $\text{H}_2\text{SO}_4$ , in the following manner:  
 $\text{HBrO}_4 + \text{H}_2\text{SO}_4 \rightarrow \text{H}_3\text{SO}_4^+ + \text{BrO}_4^-$ .

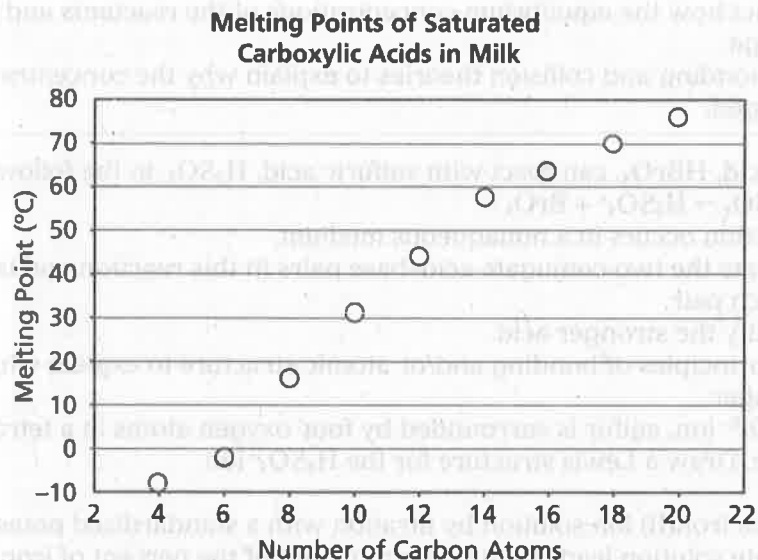
- This reaction occurs in a nonaqueous medium.
  - Indicate the two conjugate acid–base pairs in this reaction and label the acids in each pair.
  - Identify the stronger acid.
  - Use principles of bonding and/or atomic structure to explain why that acid is stronger.
- In the  $\text{SO}_4^{2-}$  ion, sulfur is surrounded by four oxygen atoms in a tetrahedral structure. Draw a Lewis structure for the  $\text{H}_3\text{SO}_4^+$  ion.

3. Analysis of an iron(II) ion solution by titration with a standardized potassium permanganate solution leads to the determination of the percent of iron in the original solution. The purple color of permanganate ion also acts as an indicator. The endpoint of the titration occurs when a very faint pink or colorless solution of manganese(II) ion is present. When performing this laboratory work, the solid of known mass and containing iron(II) is dissolved in 25.0 mL of water.

- Explain how each of the following affects the reported percentage of iron in the unknown solid:
  - The student fills the buret with  $\text{KMnO}_4$  after rinsing with only deionized water.
  - An air bubble appears in the buret tip before titration begins.
  - More than 25.0 mL of water is used to dissolve the solid.
  - Each time the volume of fluid is measured in the buret, measurement is made to the top of the meniscus, rather than the bottom of the meniscus.
- Assuming that the accepted value for the percentage of iron in the original sample was 7.77% and that your experiment gave a result of 6.896%, determine the percentage error in your work.
- Write a balanced net ionic equation for the reaction which occurs between iron(II) ions and the permanganate ion in acidic solution.

**GO ON TO NEXT PAGE**

4. The melting points of carboxylic acids found in milk are in the graph below. Saturated means all carbon-carbon bonds in the molecule are single bonds.



- (a) Identify and explain the relationship between the number of carbon atoms and the melting point of the acid in terms of the intermolecular forces present.  
 (b) Predict and explain the trend in solubility of carboxylic acids in water which makes up 87% of milk.

There are several unsaturated carboxylic acids present in milk. Unsaturated means at least one carbon-carbon double bond is present. Data for three C<sub>18</sub> carboxylic acids are listed below.

Acid Name	Number of Double Bonds	Melting Point (°C)	
Oleic	1	13	
Linoleic	2	-5	
Linolenic	3	-11	

- (c) Identify and explain the relationship between the number of double bonds and the melting point of the acid in terms of the intermolecular forces and molecular shape.
5. Explain the following observations:  
 (a) When a solid is heated at its melting point, the temperature does not increase.  
 (b) When alcohol (ethanol) is poured over your arm, your skin feels cold.

6. For a given reaction, a proposed rate law is  $\text{Rate} = k[X][Y]$ .
  - (a) Explain how the value of the rate constant,  $k$ , may be changed.
  - (b) Compare the rate of a first-order reaction with the rate of a zero-order reaction.
  - (c) Explain why knowing the rate law expression is essential when proposing a reaction mechanism.
  
7. Effervescent cold tablets are composed of 1.000 g of citric acid ( $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$ ) and 1.916 g of sodium bicarbonate. When placed into water, the two solids dissolve and react.
  - (a) Write the balanced net ionic equation for this reaction assuming that the acid is fully deprotonated in this reaction.
  - (b) Determine the theoretical volume of gas produced at STP.

In a calorimetry experiment, two tablets placed into 25.0 g of water at 21.2°C lowers the temperature of the water to 13.6°C. Assume no heat exchange with the surroundings.

- (c) The reaction is spontaneous at room temperature. Predict the sign of  $\Delta H^\circ$ ,  $\Delta S^\circ$ , and  $\Delta G^\circ$ . Justify your answers.
- (d) Calculate the energy absorbed by the water (specific heat capacity of  $\text{H}_2\text{O} = 4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1}$ ).
- (e) Calculate  $\Delta H^\circ$  for the balanced chemical equation.

END OF EXAMINATION

**ANSWERS TO DIAGNOSTIC TEST****SECTION I: MULTIPLE-CHOICE QUESTIONS**

Score your test using the table below.

Determine how many questions you answered correctly. You will find explanations of the answers on the following pages.

1. D	2. A	3. B	4. B	5. A
6. B	7. A	8. B	9. C	10. C
11. C	12. D	13. A	14. B	15. A
16. C	17. A	18. A	19. D	20. A
21. D	22. D	23. C	24. A	25. A
26. A	27. B	28. A	29. D	30. B
31. D	32. A	33. C	34. C	35. B
36. D	37. B	38. A	39. B	40. D
41. A	42. C	43. A	44. B	45. B
46. B	47. B	48. C	49. D	50. A
51. B	52. D	53. A	54. B	55. B
56. A	57. B	58. D	59. D	60. A

**CALCULATE YOUR SCORE:**

Number answered correctly: \_\_\_\_\_

**WHAT YOUR SCORE MEANS:**

Each year, since the test is different, the scoring is a little different. But generally, if you scored 20 or more on the multiple-choice questions, you'll most likely get a 3 or better on the test. If you scored 28 or more, you'll probably score a 4 or better. And if you scored 40 or more, you'll most likely get a 5. Keep in mind that the multiple-choice section is worth 50% of your final grade, and the free-response section is worth 50% of your final grade. To learn more about the scoring for the free-response questions, turn to the last page of this section.

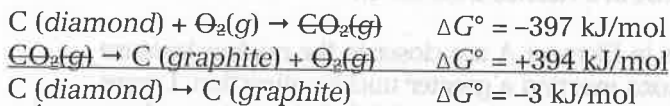
**ANSWERS AND EXPLANATIONS****SECTION I: MULTIPLE-CHOICE QUESTIONS**

- ANSWER: D** Filtration can be used to separate a heterogeneous mixture while the other techniques can be used to separate homogeneous mixtures (*Chemistry* 8th ed. pages 26–28/9th ed. pages 27–31). LO 1.19

2. **ANSWER: A** The positive slope at the left part of the graph indicates a direct relationship between the mass of calcium chloride and the mass of calcium carbonate. This means that  $\text{CaCl}_2$  is the limiting reagent and is determining the mass of  $\text{CaCO}_3$  produced (*Chemistry* 8th ed. pages 107–115/9th ed. pages 114–123). LO 3.4
3. **ANSWER: B** In the horizontal portion of the graph, the mass of  $\text{CaCl}_2$  is not related to the mass of  $\text{CaCO}_3$  produced. Therefore, the constant amount of  $\text{Na}_2\text{CO}_3$  (6.00 mL of 1.7 M solution must be responsible for the horizontal portion of the graph and must be the limiting reagent (*Chemistry* 8th ed. pages 107–115/9th ed. pages 114–123). LO 3.4
4. **ANSWER: B** The spectrum shows 3 completely filled energy levels. The peak at 347 MJ (2 electrons) is energy level 1. The 2 peaks at 29.1 and 37.1 MJ (8 electrons) is energy level 2. The 2 peaks at 2.38 and 3.93 MJ (8 electrons) is energy level 3. The last peak at 0.42 MJ is 1 electron in energy level 4. LO 1.5
5. **ANSWER: A** 0.42 MJ means that electron requires the least amount of energy to be removed from the atom. It has the lowest attraction to the nucleus, therefore it must be in a valence shell. LO 1.7
6. **ANSWER: B** The electrons in Element A are closer to the nucleus because there must be more protons exerting a greater nuclear attraction. Larger energy values indicate the electrons are more tightly held to the nucleus. LO 1.6
7. **ANSWER: A** The electron with 0.42 MJ of energy needed for its removal is a valence electron that is farther from the nucleus. It experiences less of an effective nuclear charge and can more easily be removed by water. LO 1.10
8. **ANSWER: B** Increasing the  $[\text{Ag}^+]$  will cause a shift toward products, causing an increase in the potential difference of this cell. This would, over time, cause an increase in the concentration of copper(II) ions, making the blue of the copper(II) ions darker (*Chemistry* 8th ed. pages 836–842/9th ed. pages 852–858). LO 3.12
9. **ANSWER: C** By definition the anode is the site of oxidation. Oxidation is the loss of electrons. Copper must lose two electrons to become copper(II) ions. Half-reactions involving a decrease in positive charge, like responses (B) and (D), require a gain of electrons (*Chemistry* 8th ed. pages 823–833/9th ed. pages 839–849). LO 3.13
10. **ANSWER: C** Higher temperature causes atoms in the molecule to vibrate with increased energy which leads to greater deviation from the equilibrium bond angle. LO 5.1
11. **ANSWER: C** This is a very straightforward use of the chemical formula and your understanding of moles. Since there are two moles of silver atoms per formula unit of silver sulfide, there are twice as many moles of silver atoms as moles of formula units of the

silver sulfide (*Chemistry* 8th ed. pages 81–84/9th ed. pages 85–90).  
LO 3.3

12. **ANSWER: D** From the equation  $\text{Zn}(\text{OH})_2(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2 \text{OH}^{-}(\text{aq})$  you can see that if  $2.0 \times 10^{-6}$  mol/L of the zinc hydroxide dissolves, that will result in the same concentration of zinc ions ( $2.0 \times 10^{-6}$  mol/L) and twice that much of hydroxide ions ( $2.0 \times 10^{-6}$  mol/L  $\times$  2 =  $4.0 \times 10^{-6}$  mol/L). Then:  $K_{\text{sp}} = [\text{Zn}^{2+}][\text{OH}^{-}]^2 = (2.0 \times 10^{-6})(4.0 \times 10^{-6})^2 = 3.2 \times 10^{-17}$  (*Chemistry* 8th ed. pages 744–752/9th ed. pages 759–768). LO 6.21
13. **ANSWER: A** Note that the second equation is the reverse of the first, and has been doubled. Therefore the  $K_c$  for the second equation will be the reciprocal of the first squared:  $(1/4)^2 = 1/16$  (*Chemistry* 8th ed. pages 597–601 esp. Example 13.2 and “Conclusions About the Equilibrium Expression”/9th ed. pages 610–614 esp. Example 13.2 and “Conclusions About the Equilibrium Expression”). LO 6.2
14. **ANSWER: B** To calculate  $\Delta G^\circ$  for  $\text{C}(\text{diamond}) \rightarrow \text{C}(\text{graphite})$ , you must flip the second equation and then add the equations together.



Negative values for  $\Delta G^\circ$  indicate spontaneous reactions. The high  $E_a$  makes this a kinetically slow reaction (*Chemistry* 8th ed. pages 790–794/9th ed. pages 805–810). LO 5.18

15. **ANSWER: A** Both CaO and MgO are ionically bonded. Both involve oxygen; therefore the difference in melting temperatures must be due to the difference in attraction developed by Mg compared to that of Ca. Both are 2+ ions, but magnesium ions are smaller, meaning that the charge is more concentrated in the  $\text{Mg}^{2+}$  ion. The combination of charge and size factors, called the charge density, affects the force of attraction that the ion has; ionic charge / ionic radius = charge density (*Chemistry* 8th ed. pages 318–327/9th ed. pages 329–339). LO 2.24
16. **ANSWER: C** Equilibrium tends to shift to the side with the least number of moles to relieve pressure. When the number of gaseous moles is equal, there will be no change to the system. This effect of a change in pressure on an equilibrium system is described by LeChâtelier’s principle (*Chemistry* 8th ed. pages 620–626 esp. Sample Exercise 13.14/9th ed. pages 633–639 esp. Sample Exercise 13.14). LO 6.8
17. **ANSWER: A** The reaction forms  $\text{BaSO}_4(\text{s})$  which means that ions are removed from solution, thereby making the solution less conductive. However, continued addition of the acid causes the conductivity to increase again with the excess of ions (*Chemistry* 8th ed. pages 144–150/9th ed. pages 153–158). LO 2.8



18. **ANSWER: A** The law of multiple proportions is illustrated by two elements which form at least two compounds. It then compares the ratios of the masses of one element with a constant mass of the other element in the two (or more) compounds. These ratios always reduce to simple whole numbers (*Chemistry* 8th ed. pages 41–44/9th ed. pages 44–47). LO 1.1
19. **ANSWER: D** Using the general equation  $B + HOH \rightleftharpoons HB^+ + OH^-$ ,  
 $K_b = [HB^+][OH^-]/[B]$   
 $1.0 \times 10^{-5} = x^2/0.10$   
 $x^2 = 1.0 \times 10^{-6}$   
 $[OH^-] = 1.0 \times 10^{-3} M$      $pOH = 3.00$   
 $pH = 14.00 - 3.00 = 11.00$   
 (*Chemistry* 8th ed. pages 665–666/9th ed. pages 679–680). LO 6.16
20. **ANSWER: A** Isotopes are found for every element and differ from each other in mass for the same element due only to a different number of neutrons in the nucleus (*Chemistry* 8th ed. pages 50–52, 77–80/9th ed. pages 54–55, 82–85). LO 1.14
21. **ANSWER: D** Begin by determining the formula for iron(II) phosphate:  $Fe_3(PO_4)_2$ . From this you can see that there are two phosphate ions with four oxygens in each, for a total of 8 moles of oxygen for each mole of  $Fe_3(PO_4)_2$  (*Chemistry* 8th ed. pages 56–67/9th ed. pages 59–70). LO 1.4
22. **ANSWER: D** When solving  $K_{sp}$  problems, first write the equation with the solid on the left and the ions on the right:  $MgF_2(s) \rightleftharpoons Mg^{2+}(aq) + 2 F^-(aq)$ . If  $s$  represents the mol/L of the solid that goes into solution,  $[Mg^{2+}]$  also equals  $s$ , and  $2s$  represents the  $[F^-]$ . Next write the  $K_{sp}$  expression in terms of  $s$ :  $K_{sp} = (s)(2s)^2 = 4s^3 = 6.4 \times 10^{-9}$ .  $s = \sqrt[3]{(6.4 \times 10^{-9}/4)} M$  (*Chemistry* 8th ed. pages 743–747/9th ed. pages 759–763). LO 6.21
23. **ANSWER: C** A good clue as to the type of titration curve is the position of the equivalence point, which is the center of the vertical section of the graph showing the very rapid change of pH. In this case, that is in the basic region (above 9). At this point, the amount of added  $OH^-$  equals the original amount of acid. The pH exceeds 7 (neutral) due to the hydrolysis of the anion from the acid (*Chemistry* 8th ed. pages 713–715, 717–725/9th ed. pages 727–729, 731–739). LO 6.13
24. **ANSWER: A** The leveling-off shown between F and G is caused by buffering. Optimal buffering occurs when  $[HA] = [A^-]$ , which would be at a volume of 25 mL of base in this case (*Chemistry* 8th ed. pages 710–712, 717–725/9th ed. pages 724–726, 731–739). LO 6.20

25. ANSWER: A

	<b>name</b>	<b>formula</b>	<b>IMF Present</b>
a.	propane	CH <sub>3</sub> CH <sub>2</sub> CH <sub>3</sub>	London dispersion
b.	ethanol	CH <sub>3</sub> CH <sub>2</sub> OH	London dispersion, dipole-dipole, hydrogen bonding
c.	dimethyl ether	CH <sub>3</sub> OCH <sub>3</sub>	London dispersion, dipole-dipole

The more easily a substance can attain the vapor state, the more molecules will be available to exert vapor pressure. Propane molecules only have to overcome relatively weak London dispersion forces to go into the gas phase. The other molecules have stronger IMF to overcome (*Chemistry* 8th ed. pages 440–443, 471–479/9th ed. pages 455–458, 483–491). LO 2.16

26. ANSWER: A Lone pairs of electrons (unbonded pairs) are less associated with the attractive forces of positive nuclei, therefore will occupy more space than bonded pairs (*Chemistry* 8th ed. pages 380–382/9th ed. pages 391–393). LO 2.17

27. ANSWER: B

<b>formula</b>	<b>IMF Present</b>
CH <sub>4</sub>	London dispersion
CH <sub>3</sub> Cl	London dispersion, dipole-dipole
CH <sub>3</sub> OH	London dispersion, dipole-dipole, hydrogen bonding
RbCl	Ion-ion

The temperature at which substances boil is a function of the forces between the molecules of that substance (intermolecular forces). Small, nonpolar molecules boil at the lowest temperatures since they have the weakest IMF to overcome (London dispersion). As polarity increases, so does the boiling temperature. Ionic substances have the highest attraction which means the boiling temperature is the highest (*Chemistry* 8th ed. pages 440–443/9th ed. pages 455–458). LO 2.16

28. ANSWER: A Before reaction, no combination of NH<sub>3</sub> and Ag<sup>+</sup> has formed; the silver nitrate is found as ions, the ammonia is molecular, and water is also present (*Chemistry* 8th ed. pages 759–762/9th ed. pages 774–777). LO 3.2

29. ANSWER: D The VSEPR model is very helpful in describing almost all molecular shapes. Note that in PH<sub>3</sub> the shape may, at first, seem to be tetrahedral (4 electron domains) or trigonal planar (due to the formula), but it does not have a hydrogen in one of the four apex positions, so it is a trigonal pyramid. Molecular shape is determined by the position of the nuclei and not by the position of just electrons, either pairs or single electrons (*Chemistry* 8th ed. pages 378–390, esp. Figure 8.16/9th ed. pages 389–402, esp. Figure 8.16). LO 2.21

30. **ANSWER: B** Ionic bonds between  $\text{Na}^+$  and  $\text{Cl}^-$  are broken when  $\text{NaCl}$  dissolves. Hydrogen bonds between some water molecules are broken and the ions form ion-dipole bonds with water. The ion-dipole bonds are comparable in strength to covalent bonds (*Chemistry* 8th ed. pages 130–132; 501–504/9th ed. pages 139–141, 514–517). LO 5.10

31. **ANSWER: D** The need for resonance structures seems to be greatest for molecules in which the same two elements are bonded with different type bonds (e.g., one single, one double bond) in the same molecule. The usual Lewis structure for sulfur dioxide shows that the sulfur and oxygen are bonded with one single bond and, in the other sulfur-oxygen bond, doubly bonded. (Be sure you can draw this Lewis structure.) (*Chemistry* 8th ed. pages 374–378/9th ed. pages 385–389). LO 2.21

32. **ANSWER: A**

$$K = \frac{[\text{C}]^2}{[\text{A}][\text{B}]}$$

$$K = \frac{(0.05)^2}{(0.25)(0.25)}$$

This math can be simply calculated by changing the decimals into fractions and using mathematical manipulation to put all numbers in the numerator for ease of multiplication.

$$K = \left(\frac{1}{20}\right)\left(\frac{1}{20}\right)\left(\frac{4}{1}\right)\left(\frac{4}{1}\right) = \frac{16}{400} \text{ which reduces to } 4/100 \text{ or } 0.040$$

(*Chemistry* 8th ed. pages 597–601/9th ed. pages 610–614). LO 6.5

33. **ANSWER: C**

$$K_a \times K_b = K_w$$

$$K_b = 1.0 \times 10^{-14} / 5.0 \times 10^{-10} = 2.0 \times 10^{-5} \text{ (} \textit{Chemistry} \textit{ 8th ed. pages 671–677/9th ed. pages 686–691). LO 6.16}$$

34. **ANSWER: C** While using actual electronegativities would be helpful for this question, a good generalization is that the further apart the elements are on the periodic table the greater the difference in control they have over the shared electron pair, and the more polar the bond. Note also that this is a question dealing with the polarity of an individual bond and not with the polarity of an entire molecule (*Chemistry* 8th ed. pages 344–346/9th ed. pages 356–358). LO 2.18

35. **ANSWER: B** The conjugate base of a weak acid is reacting with water (this is hydrolysis):  $\text{C}_2\text{H}_3\text{O}_2^- + \text{HOH} \rightleftharpoons \text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^-$ .  $\text{OH}^-$  raises the pH (*Chemistry* 8th ed. pages 717–725/9th ed. pages 731–739). LO 6.13

36. **ANSWER: D** Mass spectra assign the largest peak (most abundant isotope) a value of 100%, but that does not mean that isotope has 100% abundance. Assume abundance of  $^{37}\text{Cl}$  is  $x$ . Examining the relationship between the lines shows that  $^{35}\text{Cl}$  is 3 times larger than  $^{37}\text{Cl}$  or  $3x$ .  $x + 3x = 100$  since the total abundance must equal

100%. Therefore  $^{37}\text{Cl}$  is 25% and  $^{35}\text{Cl}$  is 75% (*Chemistry* 8th ed. pages 49–52, 77–80/9th ed. pages 54–55, 82–85). LO 1.14

37. **ANSWER: B** Ions with negative charges have gained an electron and are therefore larger than their parent atoms. Between the chloride ions and the bromide ion you have outermost electrons in the third energy level versus the fourth energy level, hence we would expect  $\text{Br}^-$  to be larger (*Chemistry* 8th ed. pages 349–353, esp. Figure 8.8/9th ed. pages 361–365, esp. Figure 8.8). LO 1.9
38. **ANSWER: A** Since  $\text{HBr}$  is a strong acid (completely dissociated), the  $[\text{H}^+]$  would be  $0.001\text{ M}$  which gives a  $\text{pH} = -\log 0.001 = 3.0$  (*Chemistry* 8th ed. pages 642–651/9th ed. pages 656–666). LO 6.15
39. **ANSWER: B** The substances you might expect to form solids due to low solubility might be lead(II) chromate and potassium nitrate. There are two rules which suggest the latter is not going to precipitate: “most 1A compounds are soluble,” and “most nitrate salts are soluble” (*Chemistry* 8th ed. pages 144–149/9th ed. pages 153–157). LO 6.22
40. **ANSWER: D** If gas is allowed to escape, then there are fewer molecules left within the cylinder; these fewer molecules exert less pressure (gas pressure is due to the collision of gas particles with the sides of the container). These remaining gas molecules occupy the entire container, so there must be more room between the individual molecules (*Chemistry* 8th ed. pages 181–183, 205–212/9th ed. pages 190–192, 214–222). LO 2.6
41. **ANSWER: A** Compounds with these properties are ionically bonded. This suggests elements from the far left side of the periodic table (1A or 2A) with elements from the far right side (6A or 7A) (*Chemistry* 8th ed. pages 357–358, 468–471/9th ed. pages 369–370, 480–483). LO 2.19
42. **ANSWER: C** While this may appear to be an equilibrium question, it is actually just asking you to show an understanding of the stoichiometry between ammonia and hydrogen gases as this reaction proceeds. If enough ammonia reacts to cause a decrease of  $0.40\text{ atm}$  in pressure, then enough hydrogen is formed to cause an increase in pressure of  $0.60\text{ atm}$ , ( $0.40\text{ atm} \times 3\text{ mol H}_2/2\text{ mol NH}_3$ ) (*Chemistry* 8th ed. pages 101–107/9th ed. pages 108–114). LO 2.6
43. **ANSWER: A** Go back to first principles: What is pressure? The basic kinetic-molecular theory of gases indicates that gas pressure is due to molecular collisions and more collisions mean greater pressure (*Chemistry* 8th ed. pages 205–212/9th ed. pages 214–222). LO 2.4
44. **ANSWER: B** Nonideal (real) gases differ from ideal gases because of two factors: attractions between real gas molecules and the molecular volume of real gases.  $\text{Xe}$  has 54 electrons which make the atom very polarizable so it has significant IMF. It also has the

largest atomic volume of the gases listed (*Chemistry* 8th ed. pages 214–216/9th ed. pages 224–226). LO 2.4

45. **ANSWER: B** Hydrogen bonds exist between the nitrogenous bases in DNA. This relatively strong IMF must absorb energy to overcome the attraction between the atoms of the different strands of DNA (*Chemistry* 8th ed. pages 441–442/9th ed. pages 456–457). LO 5.9

46. **ANSWER: B** This stoichiometry problem may cause you to reach for your calculator, but that tool is not needed; note the widely differing answers for the volume of the gas. 0.40 mol of  $\text{KClO}_3$  will form 0.60 mol of oxygen gas. The molar volume of gases at STP is 22.4 L/mol, so the volume is around 14 L ( $0.60 \text{ mol} \times 22 \text{ L/mol}$ ) (*Chemistry* 8th ed. pages 101–107, 194–199/9th ed. pages 108–114, 203–208). LO 3.4

47. **ANSWER: B** The part of the pressure (partial pressure) due to a gas in a mixture of gases is the same as the mole ratio of that gas to the total, which is the same as the ratio of the number of molecules of that gas to the total number of molecules. In this case there are 3 moles of gas total of which Ne is 1 part, hence  $1/3$  of the total pressure is due to the Ne (*Chemistry* 8th ed. pages 199–205/9th ed. pages 208–214). LO 2.6

48. **ANSWER: C** The atomic volume decreases due to an increase in nuclear charge; note that the “added” electrons are going into the same principal energy level as you go from left to right across the same period (*Chemistry* 8th ed. pages 318–323 esp. Figure 7.34, 908–912 esp. Figure 19.2/9th ed. pages 329–334 esp. Figure 7.34, 927–932 esp. Figure 19.2). LO 1.9

49. **ANSWER: D** Since this occurs as a one step reaction, we can write that the rate =  $k[\text{A}]^2[\text{B}]$ . Initially the rate will be proportional to  $(3)^2(2) = 18$ . Later, [A] becomes  $3 - 2 = 1$ , and [B] becomes  $2 - 1 = 1$ , so the rate is then proportional to  $(1)^2(1) = 1$ . That means a change in the rate by a factor of 18 to 1, or that the rate is now  $1/18$  as large as it was originally (*Chemistry* 8th ed. pages 547–565/9th ed. pages 559–577). LO 4.2

50. **ANSWER: A** This time you need to use the form  $k = \text{rate}/[\text{X}][\text{Y}]$  to determine the value of the rate constant,  $k$ .

$$\begin{aligned} k &= 4.0 \times 10^{-6} \text{ mol L}^{-1} \text{ min}^{-1} / (4.0 \times 10^{-1} \text{ mol L}^{-1})(4.0 \times 10^{-1} \text{ mol L}^{-1}) \\ &= 1/4 \times 10^{-4} \text{ L mol}^{-1} \text{ min}^{-1} = 0.25 \times 10^{-4} \text{ L mol}^{-1} \text{ min}^{-1} \\ &= 2.5 \times 10^{-5} \text{ L mol}^{-1} \text{ min}^{-1} \end{aligned}$$

(Note that if set up in this way, the mathematics become rather easy to handle.) Do watch units for rate constants, which, unlike equilibrium constants, traditionally have units assigned (*Chemistry* 8th ed. pages 547–565/9th ed. pages 559–577). LO 4.2

51. **ANSWER: B** There is almost always a question on kinetics like this on a test. Note that it is not just the greater number of molecular collisions that cause the reaction to occur faster but an increase in

the fraction of high-energy molecules which can then obtain the activation energy requirement and reaction that cause this rate increase (*Chemistry* 8th ed. pages 565–575, esp. Figure 12.12/9th ed. pages 577–588, esp. Figure 12.12). LO 4.6

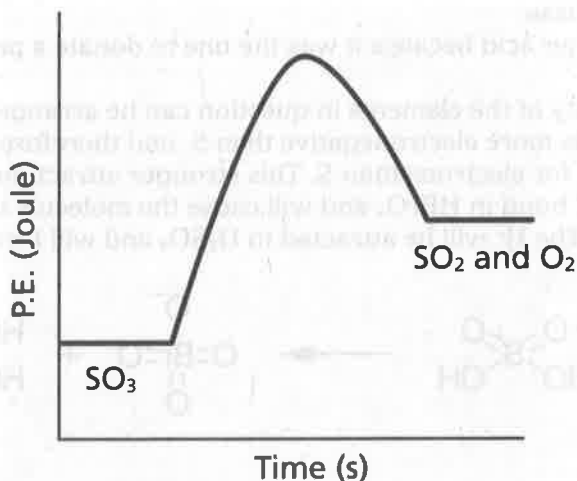
52. **ANSWER: D** Catalysts lower the activation energy barrier by forming a different activated complex (*Chemistry* 8th ed. pages 570–575, esp. Figures 12.15 and 12.16/9th ed. pages 583–588, esp. Figures 12.15 and 12.16). LO 4.8
53. **ANSWER: A** For a zero-order rate law,  $\text{rate} = k[\text{A}]^0$  (i.e.,  $n = \text{zero}$ ). Since any number taken to the zero power is equal to 1,  $\text{rate} = k$  (the rate constant). This has the physical meaning that the rate does not speed up or slow down over time (like most reactions do); it either takes place at a constant rate or does not take place at all (*Chemistry* 8th ed. pages 551–560, Table 12.6 on page 561, 576–577/9th ed. pages 563–572, Table 12.6 on page 574, 588–589). LO 4.2
54. **ANSWER: B** This is an application of the half-life expression for first-order rate laws,  $t_{1/2} = 0.693/k$ . Note that half-life (time) and the rate constant are inversely related; hence the element with the lower rate constant will have the longer half-life (*Chemistry* 8th ed. pages 551–560/9th ed. pages 563–572). LO 4.3
55. **ANSWER: B** If you wish to determine how order is affected by a given concentration, then a series of reactions should be run with changes in only the concentration of that substance, with all other concentrations held constant (as well as all other conditions like temperature held constant) (*Chemistry* 8th ed. pages 551–560 esp. Sample Exercise 12.5/9th ed. pages 563–572 esp. Sample Exercise 12.5). LO 4.1
56. **ANSWER: A** Gases have the most random organization and have the most microstates available. Gas formation is a driving force in chemical and physical processes due to the increase in entropy (*Chemistry* 8th ed. pages 773–779/9th ed. pages 788–794). LO 5.12
57. **ANSWER: B** Substances with larger IMF must absorb more energy to change phase. The longer horizontal portions of substance X indicate energy is being used to overcome IMF (*Chemistry* 8th ed. pages 471–478/9th ed. pages 483–490). LO 5.6
58. **ANSWER: D** Both substances were on the same hot plate so both were receiving the same amount of energy. The substances heated at different rates due to differences in specific heat capacity (*Chemistry* 8th ed. pages 471–478/9th ed. pages 483–490). LO 5.5
59. **ANSWER: D** This is an example of a calorimetry lab. To determine the energy gained or released ( $q$ ), you will use  $q = s \times m \times \Delta T$  where  $s$  is the specific heat capacity,  $m$  is the mass of the system, and  $\Delta T$  is the temperature change. In order to calculate for one mole, you will need volume and molarity of the solutions (*Chemistry* 8th ed. pages 243–251/9th ed. pages 252–260). LO 5.7

60. ANSWER: A This is a reaction between a strong acid and a strong base and it is exothermic so the temperature will increase. The increase is due to the large amount of energy released when  $H^+$  and  $OH^-$  bond to form  $H_2O$  (Chemistry 8th ed. pages 243–251/9th ed. pages 252–260). LO 5.8

## SECTION II: FREE-RESPONSE QUESTIONS

### Question 1: Answers

(a)



$$(b) K_C = \frac{[O_2][SO_2]^2}{[SO_3]^2}$$

This question could have asked you to calculate  $K_p$  instead. If that were the case, partial pressures would be shown in parentheses and the equation would have been

$$K_p = \frac{(O_2)(SO_2)^2}{(SO_3)^2}$$

$$(c) [SO_2] = 1.23 \times 10^{-2} \text{ mol}/2.25 \text{ L} = 0.00547 \text{ M}$$

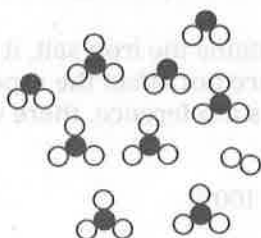
$$[O_2] = (\text{one-half as much as above}) = 0.00273 \text{ M}$$

$$\text{initial mole } SO_3 = 3.21 \text{ g}/80.1 \text{ g/mol} = 0.0401 \text{ mol}$$

$$[SO_3] = (0.0401 \text{ mol} - 0.0123 \text{ mol})/2.25 \text{ L} = 0.0124 \text{ M}$$

$$(d) K_C = (0.00273)(0.00547)^2/(0.0124)^2 = 5.31 \times 10^{-4}$$
 Recall that units are not shown on equilibrium constants.

(e) To represent the equilibrium mixture, one model will represent approximately 0.002 M. 3 particles of  $SO_2$ , 1 particle of  $O_2$ , and 6 particles of  $SO_3$ .



(f) (i)  $[SO_3]$  will decrease and  $[SO_2]$  and  $[O_2]$  will increase.

- (ii) Since the reaction is endothermic, a higher temperature will force the reaction to products to relieve the stress according to Le Chatelier's principle. The higher temperature means that more particles will have sufficient energy to overcome the activation energy and produce products.

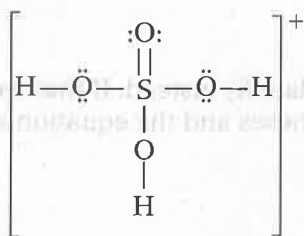
(Chemistry 8th ed. pages 602–604, 609–614/9th ed. pages 615–616, 621–627)

### Question 2: Answers

- (a) (i) The pairs are  $\text{HBrO}_4$  (acid)/ $\text{BrO}_4^-$  (conjugate base) and  $\text{H}_2\text{SO}_4$  (base)/ $\text{H}_3\text{SO}_4^+$  (conjugate acid). Note that acid form always has one more  $\text{H}^+$  than its conjugate base.
- (ii)  $\text{HBrO}_4$  is the stronger acid because it was the one to donate a proton to  $\text{H}_2\text{SO}_4$ .
- (iii) The electronegativity of the elements in question can be arranged as  $\text{O} > \text{Br} > \text{S} > \text{H}$ . Br is more electronegative than S, and therefore Br has a stronger attraction for electrons than S. This stronger attraction increases the polarity of the O–H bond in  $\text{HBrO}_4$  and will cause the molecule to dissociate into  $\text{H}^+$  and  $\text{BrO}_4^-$ . The  $\text{H}^+$  will be attracted to  $\text{H}_2\text{SO}_4$  and will form the  $\text{H}_3\text{SO}_4^+$  ion.



(b)



(Chemistry 8th ed. pages 640–642, 645, 679–682/9th ed. pages 654–656, 659, 693–695)

### Question 3: Answers

- (a) (i) The  $\text{KMnO}_4$  is actually less concentrated than you believe it to be since it has been diluted with the water left in the buret. That means that more of the standardized solution will be needed to react with the iron(II) solution leading you to report that the iron concentration is higher than it actually is.
- (ii) If the air bubble stays in the buret during the entire titration, then there is no effect. However, if it is replaced by the  $\text{KMnO}_4$  solution during the titration, then less solution leaves the buret than you are reporting, leading to a higher than correct percentage of the iron(II).
- (iii) If more water is added to the flask that contains the iron salt, it will not affect the number of moles of iron(II) and therefore not affect the reported result.
- (iv) If the same point on the meniscus is used as a reference, there will not be any effect on the percentage of iron reported.

$$(b) \quad \% \text{ Error} = \frac{(\text{your lab results} - \text{accepted value})}{\text{accepted value}} \times 100\%$$

$$\frac{(6.896 - 7.77)}{7.77} \times 100\% = -11\%$$

Note that the answer is known only to two significant figures.



(c)  $5 \text{Fe}^{2+}(\text{aq}) + \text{MnO}_4^{-}(\text{aq}) + 8 \text{H}^{+}(\text{aq}) \rightarrow 5 \text{Fe}^{3+}(\text{aq}) + \text{Mn}^{2+}(\text{aq}) + 4 \text{H}_2\text{O}(\text{l})$ .  
 (Chemistry 8th ed. pages 156–158, 165–168/9th ed. pages 165–169, 175–177)

#### Question 4: Answers

- (a) Carboxylic acids have a nonpolar chain that has London dispersion forces as the main IMF and the functional group has a polar C=O (carbonyl group) that has dipole–dipole forces and an O–H (hydroxyl group) that can form hydrogen bonds. When the number of carbons is low, the polar end of the molecule with dipole–dipole and hydrogen bonding predominates. As the number of carbon atoms increases, the proportion of London dispersion forces also increases. As the strength of the London dispersion forces increases, the molecules will experience greater attractions. More energy is needed to overcome these attractions so the melting point is higher.
- (b) The solubility of the carboxylic acids in water decreases as the number of carbon atoms increases. The predominant IMF in water is hydrogen bonding. Water is able to interact with other molecules whose predominant IMFs are dipole–dipole or hydrogen bonding. As the number of carbon atoms increases, the London dispersion forces become predominant and water cannot interact with those molecules.
- (c) As the number of double bonds increases, the melting point of unsaturated carboxylic acid decreases. Even though the molecules are approximately equal in terms of number of electrons and therefore in terms of London dispersion forces, the shape of the molecules limits the locations of interactions between the molecules. The mono-unsaturated oleic acid has a greater chance of stacking and having many locations where electrons from each molecule can interact. The di- and tri-unsaturated acids have shapes that do not allow close packing and therefore they have fewer interactions and lower melting points. The more interactions molecules can form, the more energy must be added to overcome attractions, which leads to higher melting points.

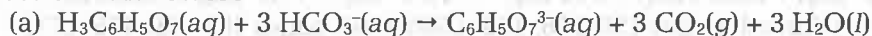
#### Question 5: Answers

- (a) Temperature is directly related to average kinetic energy. At the melting point, the thermal energy being added to the system is overcoming the attractive forces (potential energy) between the particles. Once all of the attractive forces (IMF) are broken, the temperature increases due to an increase in the average kinetic energy of the particles.
- (b) An endothermic process causes this. The alcohol evaporates. This requires energy which is lost by your skin, so the temperature of your skin decreases.  
 (Chemistry 8th ed. pages 238–244, 475–477/9th ed. pages 248–253, 487–489)

#### Question 6: Answers

- (a) The rate constant may be changed by changing the temperature (but not by changing the concentration of X or of Y). Recall the Arrhenius equation.
- (b) For zero-order reactions, the rate is constant {Rate =  $k[\text{A}]^0 = k(1) = k$ }. This means that the rate does not change with concentration as it does with a first-order reaction {Rate =  $k[\text{A}]^1$ }.
- (c) All reactants in the rate determining step must be part of the rate law expression because each of those reactants controls the reaction rate.  
 (Chemistry 8th ed. pages 547–565, Table 12.6 on page 562/9th ed. pages 559–577, Table 12.6 on page 574)

## Question 7: Answers



(b)  $\frac{1.000 \text{ g H}_3\text{C}_6\text{H}_5\text{O}_7}{192.12 \text{ g/mol}} = 0.005205 \text{ mol}$

$\frac{1.916 \text{ g NaHCO}_3}{84.01 \text{ g/mol}} = 0.02281 \text{ mol}$

	$\text{H}_3\text{C}_6\text{H}_5\text{O}_7 +$	$3 \text{HCO}_3^- \rightarrow$	$\text{C}_6\text{H}_5\text{O}_7^{3-} +$	$3 \text{CO}_2 +$	$3 \text{H}_2\text{O}$
initial mol	0.005205	0.02281	0	0	0
change mol	-x	-3x	+x	+3x	+3x
final mol	0	0.0072	0.005205	0.01561	0.01561

Since one of the reactants must be used up  $x = 0.005205$  or  $x = 0.02281/3 = 0.007603$ .  $x \neq 0.007603$  because neither reactant can have a negative number of moles at the end of the reaction. Therefore  $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$  is the limiting reactant.

$0.01561 \text{ mol CO}_2 \times 22.414 \text{ L mol}^{-1} = 0.3499 \text{ L}$

(c)  $\Delta H^\circ = +$ ; the reaction is endothermic since the temperature of the water decreased; energy flowed from surroundings into system $\Delta S^\circ = +$ ; the reaction began with 2 solids (or aqueous solutions) and formed a gas which has more microstates than other states of matter $\Delta G^\circ = -$ ; since the reaction was spontaneous,  $\Delta G^\circ$  must be negative (definition of a spontaneous process)

(d)  $q = m \times c \times \Delta T$

$= (25.0 \text{ g})(4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1})(-7.6^\circ\text{C}) = -790 \text{ J}$

(e)  $q_{\text{H}_2\text{O}}$ , 2 tablets after each  $q = -790 \text{ J}$  so  $q_{\text{2 tablets}} = +790 \text{ J}$ . 2 tablets =  $2 \times 0.005205 \text{ mol H}_3\text{C}_6\text{H}_5\text{O}_7 = 0.01041 \text{ mol}$  $790 \text{ J}/0.01041 \text{ mol} = 76000 \text{ J/mol} = 76 \text{ kJ/mol}$  (Note: This was real data collected by students; the assumption about no interaction with surroundings may not be valid.)

## SCORING THE FREE-RESPONSE QUESTIONS

It is difficult to come up with an exact score for this section of test. However, if you compare your answers to the answers in this book, remembering that each part of the test you answer correctly is worth points even if the other parts of the answer are incorrect (see the section titled "Types of Free-Response Questions" on page 12 of this book), you can get a general idea of the percentage of the questions for which you would get credit. If you believe that you got at least one-third of the possible credit, you would probably receive a 3 on this part of the test. If you believe that you would receive close to half or more of the available credit, your score would more likely be a 4 or a 5.