

Eighth Edition

Instructor's Resource Guide

Chemistry

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PART I: SCHEDULING AND LABORATORY

No two instructors will likely teach an identical course, even if they are using the same text and the same general outline. Instead of presenting several detailed schedules, we will list the minimum number of lecture periods (50 min/lecture) we feel is necessary to sufficiently cover each topic. Although the sum of lectures is 90 (three lectures/week for one year), such a course (presented below) would be too fast-paced and too sparse on the many details that flesh out topics to be practical. There really is more in the text than can be covered completely in a single year. Individual instructors need to have a basic outline and choose from the many alternatives the text offers. In the later discussion of each chapter, alternative placement and opportunities for introducing material from other chapters will be pointed out.

Listed for each topic are the numbers of appropriate experiments from the accompanying laboratory manual, *Experimental Chemistry*, Eighth Edition, by James F. Hall.

Topic	Chapter	Number of Lectures	Experiment(s)
Chemical Foundations	1	2	1, 2, 3
Atoms, Molecules, and Ions	2	3	7, 8, 9, 10, 25
Stoichiometry	3	3	11, 12, 13, 14, 15, 58
Reactions in Solution	4	5	6, 12, 13, 15, 27, 36, 37, 38, 39, 45, 46, 47, qual
Gases	5	4	16, 17, 18, 19
Thermochemistry	6	3	20
Atomic Structure and Periodicity	7	6	21, 22, 25
Bonding	8, 9	7	23, 24
Solids and Liquids	10	4	4, 5

Topic	Chapter	Number of Lectures	Experiment(s)
Properties of Solutions	11	4	27, 28, 29
Kinetics	12	4	30, 53
Equilibrium	13	4	31, 32, 33
Acids/Bases	14, 15	8	35, 36, 37, qual
Solubility	16	2	34, qual
Complex Ions	16	1	44, qual
Entropy, Free Energy, and Spontaneity	17	4	--
Electrochemistry	18	4	39, 40, 41
Nuclear Chemistry	19	4	--
Representative Elements	20	6	25, 45, 46, 47, 54, 55
Transition Elements and Coordination Chemistry	21	3	44, 45, 46, 47
Organic Chemistry	22	3	48, 49, 50, 51
Biochemistry	22	3	9, 10, 52, 53, 54, 55

An alternative to the classical, by group, qualitative analysis experiments (56-61) in the Hall manual is a more open-minded approach in *Solving Equilibrium Problems with Applications to Qualitative Analysis*, by Steven S. Zumdahl. In this extended experiment students react several reagents with 11 cations and observe what happens. They then use this database to develop their own schemes to separate unknowns containing up to 5 of the 11 cations. It is a more challenging approach to qualitative analysis. In our experience we have had the most success and positive feedback when it is used in courses composed primarily of chemical sciences majors.

EXPERIMENTAL CHEMISTRY

Eighth Edition
by James F. Hall

Introduction

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Experiments

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2. The Use of Volumetric Glassware
3. Density Determinations
4. The Determination of Boiling Point
5. The Determination of Melting Point
6. The Solubility of a Salt
7. Identification of a Substance
8. Resolution of Mixtures 1: Filtration and Distillation
9. Resolution of Mixtures 2: Paper Chromatography
10. Resolution of Mixtures 3: Thin-Layer Chromatography
11. Counting by Weighing
12. Stoichiometry 1: Limiting Reactant
13. Stoichiometry 2: Spectrophotometric Determination of the Stoichiometry of an Iron(III)-Phenol Reaction
14. Composition I: Percentage Composition and Empirical Formula of Magnesium Oxide
15. Composition 2: Percentage Water in a Hydrate
16. Preparation and Properties of Hydrogen and Oxygen Gases
17. Gas Laws 1: Charles's Law and Absolute Zero
18. Gas Laws 2: Graham's Law
19. Gas Laws 3: Molar Mass of a Volatile Liquid
20. Calorimetry
21. Spectroscopy 1: Spectra of Atomic Hydrogen and Nitrogen
22. Spectroscopy 2: Emission Lines of Some Metallic Elements
23. Molecular Properties 1: Molecular Shapes and Structures
24. Molecular Properties 2: Determination of the Length of a Molecule
25. Properties of Some Representative Elements
26. Classification of Chemical Reactions
27. Determination of Chloride Ion and Calcium Ion in Water Samples
28. Colligative Properties 1: Freezing Point Depression and the Determination of Molar Mass
29. Colligative Properties 2: Osmosis and Dialysis
30. Rates of Chemical Reactions
31. Chemical Equilibrium 1: Titrimetric Determination of an Equilibrium Constant

32. Chemical Equilibrium 2: Spectrophotometric Determination of an Equilibrium Constant
33. Chemical Equilibrium 1: Stresses Applied to Equilibrium Systems
34. The Solubility Product of Silver Acetate
35. Acids, Bases, and Buffer Systems
36. Acid-Base Titrations 1: Analysis of an Unknown Acid Sample
37. Acid-Base Titrations 2: Evaluation of Commercial Antacid Tablets
38. The Determination of Calcium in Calcium Supplements
39. Determination of Iron by Redox Titration
40. Electrochemistry 1: Chemical Cells
41. Electrochemistry 2: Electrolysis
42. Gravimetric Analysis 1: Gravimetric Determination of Chloride Ion
43. Gravimetric Analysis 1: Gravimetric Determination of Sulfate Ion
44. Preparation of a Coordination Complex of Copper (II)
45. Inorganic Preparations 1: Preparation of Sodium Thiosulfate Pentahydrate
46. Inorganic Preparations 2: Preparation of Copper(II) Sulfate Pentahydrate
47. Inorganic Preparations 3: Preparation of Sodium Hydrogen Carbonate
48. Qualitative Analysis of Organic Compounds
49. Organic Chemical Compounds
50. Ester Derivatives of Salicylic Acid
51. Preparation of Fragrant Esters
52. Proteins
53. Enzymes
54. Polymeric Substances 1: Amorphous Sulfur
55. Polymeric Substances 2: Preparation of Nylon
- 56-61. Qualitative Analysis: Techniques
56. Qualitative Analysis of the Group I Cations
57. Qualitative Analysis of the Group II Cations
58. Qualitative Analysis of the Group III Cations
59. Qualitative Analysis of the Group IV and V Cations
60. Qualitative Analysis of Selected Anions
61. Identification of an Unknown Salt

- | | |
|-------------|--|
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PART II: CHAPTER DISCUSSIONS

COURSE CONTENT: DESCRIPTIVE CHEMISTRY AND CHEMICAL PRINCIPLES

There has been much discussion and an overall agreement that more descriptive chemistry needs to be included in the general chemistry course. This has been one of the major considerations in the recent curriculum reform movement. How to accomplish this goal has resulted in much less agreement. General chemistry has to move beyond a rote presentation of the facts and principles of chemistry. We feel that the facts (descriptive chemistry) and principles must be an integrated whole that is aimed at fulfilling the basic goals of the course.

The *Chemistry* textbook follows what has been for the past several years, a fairly traditional order of topics. The chapters on principles (1-18) are followed by nuclear chemistry (19), a more descriptive chapter on chemistry of the elements (20), a chapter on transition metals and coordination chemistry (21), and organic chemistry and biochemistry (22). Different instructors will have different approaches to descriptive chemistry and, in fact, will probably give greatly different definitions of descriptive chemistry. We have organized the text this way because we feel this order offers instructors the greatest flexibility.

In the fifties and early sixties, general chemistry texts essentially went through the periodic table; principles were introduced as needed. Observations were made of the properties and reactions of matter; then we used chemical principles to try to systematically organize and understand those observations. This approach may closely resemble how chemists do chemistry. But is it the best way to learn chemistry? We don't think so. We are all driven by curiosity as to how things work, and an unfortunate consequence of marching through the periodic table is to dull the students' curiosity about why things happen, replacing it with the impression of chemistry as a vast body of unrelated facts and reactions that must be memorized. This approach tends to make learning chemistry both frustrating and uninteresting. In the sixties, a principles-dominated outline began to be used in texts, but we overcompensated. The principles were emphasized. The framework was put in place to deal with chemical facts, but all too often the facts were left out. Only the reaction $A + B \rightarrow C$ seems to have been covered. Learning chemistry was still frustrating, but in a different way from before.

We must integrate the two approaches. In truth, the principles of chemistry and descriptive chemistry are two indispensable parts of the whole. In particular, we must use all of our resources in presenting reactions. Can we honestly expect a neophyte to chemistry to appreciate the difference between



if they are only written on the blackboard or in a book? We must make use of all of the senses. Students can appreciate what $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ means if the instructor ignites a mixture of hydrogen and oxygen. They can see, hear, and sometimes feel what that equation represents.

We can't just talk about reactions. We must show the students reactions or, better yet, let them see the reactions on their own in the laboratory. We should choose more experiments for the labs in which students observe reactions or synthesize compounds and do fewer experiments that involve a measurement that confirms (often poorly) some principles discussed in class. The use of lecture demonstrations also allows students to see chemistry. These two approaches represent powerful tools for presenting reactions.

In the text we have tried to emphasize the framework of principles by which chemical facts can be organized. There are several important features that facilitate the integration of fact and principle.

1. To assist instructors in providing interesting and relevant classroom demonstrations of chemical phenomena for their students, the *Instructor's Annotated Edition* indicates via the marginal icon (flask with stirring rod) direct references to more than 750 demonstrations from several authoritative sources. The sources are provided in a separate booklet, the *Lecture Demonstration Guide* by Fred Juergens of the University of Wisconsin, Madison, that is available to adopters of the eighth edition.
2. Chapter 4 presents an early, thorough discussion of solution reactions, using the classifications that chemists typically use (acid-base, precipitation, and oxidation-reduction).
3. The illustrations and color photography show the student many reactions. This merely supplements lecture demonstrations and labs but right in the text students can see the beauty of the reactions and substances composing this body of knowledge called chemistry.
4. Real examples are used to illustrate chemical principles.
5. Much of the discussion in the descriptive chapters is organized according to a framework of the early chapters. For example, Chapters 19-21 emphasize periodic relationships; the discussion of polymers in Chapter 22 emphasizes structure-property relationships.
6. Many of the end-of-chapter exercises deal with reactions or substances that are important to our lives. In the problem statement, these relationships are mentioned.

7. The end-of-chapter exercises in Chapters 19-22 are designed to illustrate the relationships between the descriptive material and the principles chapters. Energy relationships (particularly bond energies), periodicity, structure, and equilibrium are emphasized.

In the following chapter discussions, we will further point out how descriptive material from the later chapters can be brought into discussion of principles. The organization of the text allows each instructor greater flexibility in presenting the material. The unique features of the text allow for a more thorough integration of principles and facts. Chemistry isn't just principles; it isn't just reactions and properties of elements and compounds. It is an amalgam of both. In the text we have tried to bring both together, while allowing instructors maximum flexibility.

CHAPTER ONE: CHEMICAL FOUNDATIONS

Chapter Learning Goals:

- | | |
|----------------|---|
| Section One: | To appreciate the importance of creative problem solving. |
| Section Two: | To identify the principal operations and limitations of the scientific method. |
| Section Three: | To describe the SI system of units and prefixes. |
| Section Four: | To identify causes of uncertainty in measurement.
To show how significant figures are used.
To compare precision and accuracy in measurement. |
| Section Five: | To show how to determine the number of significant figures in a calculated result. |
| Section Six: | To show how to convert units between the English and metric systems. |
| Section Seven: | To demonstrate conversions among the Fahrenheit, Celsius, and Kelvin temperature scales. |
| Section Eight: | To illustrate calculations involving density. |
| Section Nine: | To show how matter can be classified into subgroups. |

The time spent on the first two chapters may show the greatest differences from class to class. Much of the first chapter is probably review for students with good high school math and chemistry backgrounds.

This chapter lays the foundation for dealing with measured quantities and performing calculations. Instructors wishing to treat uncertainty in greater detail can discuss the section Uncertainties in Measurements, from Appendix One. This section of the appendix could also be used in the laboratory.

Instructors should be careful to point out that the Sample Exercises are worked in a different manner from what students should do. Intermediate answers are rounded off

to show the correct number of significant figures at each stage of the calculation. These rounded values are then used to complete the calculation. Students, when working problems, should round only at the end. In the *Solutions Guide*, we have followed the same convention as the text and have rounded at the points where intermediate answers are shown. If this would result in excessive round-off error, we have carried extra digits and noted this in the solution. Most commonly we have carried extra digits when solving two simultaneous equations and in doing equilibrium problems.

One of the primary reasons for the discussion of units is to introduce the student to the use of dimensional analysis as a problem-solving technique. In our classes we emphasize using units as a check. Some students may tend to be sloppy about units in calculations because they are familiar with the quantities being used. The nice feature of dimensional analysis is that it works well even if we don't have a good intuitive feel for the quantities and units encountered in a particular problem. On the other hand, it is important to remember that correct use of dimensional analysis does not mean that a student has a conceptual understanding of the chemical concepts in a particular problem. One of the reasons we have included the Active Learning Questions is so the students have a chance to vocalize their ideas about the concepts covered, and we have a better chance to try to understand these.

Teaching Tips

- Connecting the real world to the atomic/molecular world is difficult for many students. Encourage students to visualize the atomic/molecular world by drawing pictures. The student responses will give you an idea of their entering perceptions. Don't worry about some of the details such as subatomic particles here. For example, point out the representations of hydrogen, oxygen, and water in Section 1.1.
- Consider spending some time discussing Figure 1.4. Some students believe that theories become laws once a theory has been accepted, but this figure clearly shows laws and theories as two separate entities. Theories never become laws. Laws tell us what happens, and theories are our attempts to explain why. In terms of a "scientific method" students should understand that much of scientific thinking is logical or analytical thinking.
- To illustrate uncertainty in measurements, bring to class several pieces of glassware used to measure volume. Show the students the glassware. Draw the shape of a representative piece of the glassware on the board marking the drawing to show the meniscus of a liquid between two of the lines. Since the volume falls between the lines there is uncertainty involved in the reading. Have the students write down the result of the measurement. Then compare results. The number recorded for this measurement will vary from student to student only in the last digit. Thus there is only one uncertain digit.

- A relatively easy way to determine the number of significant figures in a number is to write the number in scientific notation. In this notation there are no leading zeros, all zeros in the middle of the number are significant, and trailing zeros are only recorded if they are significant.
- To illustrate the difference between a physical and chemical change, compare representation (b) and (c) from Figure 1.11 (a physical change from liquid water to gaseous water) to the chemical change when liquid water is electrolyzed to produce gases H_2 and O_2 . Use the equation in the text to represent the electrolysis of water chemical change.

CHAPTER TWO: ATOMS, MOLECULES, AND IONS

Chapter Learning Goals:

- | | |
|----------------|---|
| Section One: | To give a brief account of early chemical discoveries. |
| Section Two: | To describe and illustrate the laws of conservation of mass, definite proportion, and multiple proportions. |
| Section Three: | To describe Dalton's theory of atoms and show the significance of Gay-Lussac's experiments. |
| Section Four: | To summarize the experiments that characterize the structure of the atom. |
| Section Five: | To describe features of subatomic particles. |
| Section Six: | To introduce basic ideas of bonding in molecules.
To show various ways of representing molecules. |
| Section Seven: | To introduce various features of the periodic table. |
| Section Eight: | To demonstrate how to name compounds given their formulas and to write formulas given their names. |

Chapter 2 is the second of the two introductory chapters. As with Chapter 1, the length of time spent on this chapter will depend on the background of the students. It goes over the background in chemical topics that provide the foundation for the rest of the course.

This chapter takes a historical approach to the development of chemistry. It begins with a discussion of the discoveries leading to Dalton's atomic theory, continues with the experiments elucidating the structure of the atom, and ends with chemical nomenclature and an introduction to the periodic table.

The section on nomenclature and formula writing in the laboratory manual by Hall can be effectively used at this point. Instructors may also wish to discuss nuclear decay in more detail at this time. Chapter 19, Nuclear Chemistry, can be covered here if instructors prefer. However, if the chapter is presented early, the sections on decay kinetics and thermodynamic stability of the nucleus would have to be delayed and could

be covered with Chemical Kinetics (Chapter 12) or Spontaneity, Entropy, and Free Energy (Chapter 17).

Note that the eighth edition places an increased emphasis on pictorial representations of atoms and molecules. This emphasis shows up in many of the following chapters.

Teaching Tips

- Consider using an analogy between the letters in the alphabet and words and the atoms of the various elements and compounds. In the same way we can make many words from 26 letters, millions of compounds can be made using only the 100 or so elements. Putting the same elements together in different ways results in different compounds with very different properties. For example, the words “ADD” and “DAD” each consist of the same letters. However, the order of the letters gives them different meanings. This discussion will help later (for example, in Chapter 3 you can use this analogy to help the students understand why we do not change subscripts in a chemical formula).
- Students often have difficulty understanding what is meant by chemical terms such as element. It is important to emphasize the various ways we use the term element in chemistry. When we say element we might mean a single atom, we might mean a molecule such as N or O or we might mean a large sample such as a bar of aluminum or the graphite (“lead”) in a pencil. Be sure to help students realize that chemists use terms in many ways and they need to look at the context of the word to be sure that they understand what is intended. Chemists are accustomed to thinking about things at the macroscopic and microscopic level simultaneously. Encourage students to think about what they read and to consider how the terms are being used.
- Sections 2.2 and 2.3 provide another opportunity to discuss the difference between laws and theories. Dalton’s atomic theory is relatively simple in scope, and this is as it should be. John Dalton was trying to explain laws such as the law of constant composition. The success of this makes the model successful, but not absolutely correct (a model is always a simplification). For example, Dalton’s theory does not explain questions such as “Why/how do atoms stick together to form molecules?”, and “Why/how do molecules stick together to form liquids and solids?”. But no model answers all questions.
- The discovery of the proton leads to an excellent example of how and why models change. Dalton’s model of the atom did not account for isotopes since he assumed that all atoms of an element are exactly alike. As knowledge expands models change to accommodate the new information.
- Students should learn the names of the common groups on the periodic table (see Figure 2.19). This will simplify your discussion later in the course. Reasons behind the structure of the periodic table and trends it shows are not introduced

until Chapter 7. Focus the students' attention of the location of groups of elements and the separation between metals and nonmetals.

- Emphasize that ions are formed by the gain and loss of electrons. This is difficult for students because they tend to think about gain and loss using positive numbers. It is very important that they understand from the beginning that a positive ion is formed by losing electrons and a negative ion is formed by gaining electrons. In chemistry the positive number (protons) remain the same.
- Be sure to point out to students that atoms do not spontaneously gain or lose electrons. The gain and loss of electrons are always paired. One atom loses electrons simultaneously with another atom gaining electrons.
- Stress that learning to name compounds does not consist of memorizing a seemingly endless list of chemicals. There are systematic rules for naming compounds and, by knowing only a few rules, the students can name most any compound they will encounter in this course.
- The students should also understand that we keep the rules as simple as possible (much the same way we keep our scientific models as simple as possible). We only add complications (such as prefixes and Roman numerals) when it is required for clarity. For example, the name sodium(I) chloride is not necessarily wrong, merely redundant. Therefore, the name sodium chloride suffices.
- The students can determine the charges of all the ions in Table 2.3 from their positions on the periodic table, with the exception of Ag^+ . The students should appreciate that the periodic table contains a great deal of information. It is a valuable resource.
- Make a clear distinction between Type I and Type II cations. This reinforces the reasons for the Roman numerals in the names of the compounds. Ask students, for example "How is 'iron oxide' different from 'magnesium oxide'?"
- Students often make the mistake that the Roman numeral tells us that number of ions present in the compound. This is because in many cases the Roman numeral is the same as one of the subscripts. Use an example such as iron(II) oxide with the formula FeO to show that the Roman numeral relates only to the charge.
- For Type III compounds, consider starting with examples such as "NO" and "NO₂." Ask the students to name these compounds. This will lead naturally to a discussion of prefixes. Like Roman numerals, prefixes are used only for clarity. For example, there are two possible "carbon oxides" and we must use prefixes to

differentiate between these (carbon monoxide and carbon dioxide). We only make the rules more complicated when it is required.

- Figure 2.22 and Figure 2.23 are convenient flow charts to help students systematically name compounds. Students need to be encouraged to use this type of device to help them think through a difficult problem.
- Students should think of polyatomic ions as a “unit.” Many students know that if NaCl is dissolved in water, Na^+ ions and Cl^- ions are present. However, students are often confused about the ions present when NaNO_3 is dissolved. Instead of the NO_3^- ion students often think individual nitrogen and oxygen ions are present.

CHAPTER THREE: STOICHIOMETRY

Chapter Learning Goals:

- | | |
|-----------------|--|
| Section One: | To describe the modern atomic mass scale and explain how atomic masses are determined experimentally. |
| Section Two: | To explain atomic mass and its experimental determination. |
| Section Three: | To explain the importance of the mole concept.
To show how to convert among moles, mass, and number of particles for a given sample. |
| Section Four: | To show how to calculate values for molar mass.
To show how to convert among molar mass, moles, and number of particles in a given sample. |
| Section Five: | To describe a conceptual problem solving approach to chemistry. |
| Section Six: | To demonstrate the calculation of the mass percent of a given element in a compound. |
| Section Seven: | To demonstrate the calculation of the empirical formula of a compound.
To show how to obtain the molecular formula, given the empirical formula and the molar mass. |
| Section Eight: | To identify the characteristics of a chemical reaction and the information given by a chemical equation. |
| Section Nine: | To show how to write a balanced equation to describe a chemical reaction. |
| Section Ten: | To show how to calculate the masses of reactants and products using the chemical equation. |
| Section Eleven: | To show how to recognize the limiting reactant.
To demonstrate the use of the limiting reactant in stoichiometric calculations. |

Chapter 3 deals with the fundamental measurement unit in chemistry, the chemical mole. The law of conservation of mass is the unifying principle of the chapter. The chapter deals with compounds first and then reactions. In the eighth edition, more chemical equations are represented pictorially, consistent with the representation of atoms and molecules in the second chapter.

An important point that can be made in lecture is the convenience of the mole as a unit. Chemical reactions, chemical formulas, and structures of molecules focus on numbers of atoms, molecules, or ions. Atoms are so small that we cannot see them to count them. Even if we could see them, the numbers that we would encounter are so large it would take an immense period of time to count them. The mole provides a unit that allows us to connect the number of atoms (what we are interested in knowing) to something we can measure (such as mass). This approach makes it easy to introduce molarity in Chapter 4 as a similar type of unit. Molarity enables us to convert between the quantity we need to deal with in reactions (numbers of moles), and the most convenient measurement we can make for the amount of a solution (volume).

Teaching Tips

- Section 3.5 discusses a conceptual problem solving strategy for the students. It offers suggestions on thinking about the problem so that they learn how to solve problems in general, not simply use the right algorithm for a given problem. It is well worthwhile to spend some time with the students discussing this.
- The concept of counting atoms from the mass of a sample can be difficult for students to understand. Develop the idea slowly. Many students learn to work the problems without really understanding this section. This leads to trouble later when they try to solve more complex problems. Return to the candy analogy to assist in developing this concept. You can also use a hardware store analogy of buying nails by weight, or a banking analogy of counting coins into rolls by weighing them.
- The idea that different samples with identical mass ratios contain the same number of objects is a difficult one for the students. Consider using the following example.

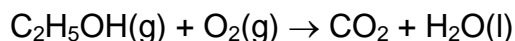
Suppose we have two blocks, a red block and a yellow block. The red block weighs 1.0 ounce, and the yellow block weighs 4.0 ounces. Now suppose we have 16 of each color block. What is the mass of each sample? The sample of red blocks weighs 16.0 ounces and the sample of yellow blocks weighs 64.0 ounces. But note that 16.0 ounces is also 1.0 pound, thus 64.0 ounces is 4.0 pounds. The relative masses of the blocks are:

	<u>one block</u>	<u>sixteen blocks</u>
red	1.0 ounce	1.0 pound
yellow	4.0 ounces	4.0 pounds

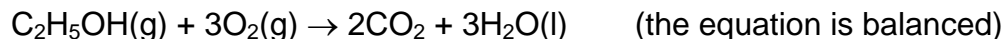
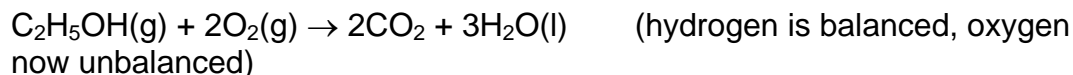
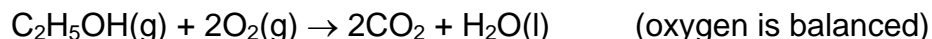
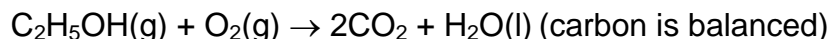
The relative masses stay the same (1:4) but the units are changed.

You can make an analogy between ounces and pounds in this case and amu's and grams in the case of the periodic table. In the blocks example, the number 16 is analogous to Avogadro's number.

- It is important for students to realize that a mole really describes the number of objects present, just as a dozen means 12. This connection between a mole as a number and the mass of a sample is essential because we count atoms by weighing samples containing large numbers of them.
- Notice that the text defines the mole differently from the SI definition: "the mole is the amount of substance that contains as many entities as there are in exactly 0.012 kg of carbon 12." The text definition is "the mole is the number equal to the number of carbon atoms in 12.01 grams of carbon." We believe the text definition is easier for the students to understand at this point as it more strongly emphasizes both that the mole is fundamentally a number and that the periodic table contains average atomic masses.
- In Section 3.9 students learn that a chemical reaction involves rearrangement of the elements. It is very helpful to use simple models of the reactants that you can take apart and rearrange to show students that atoms are conserved in a chemical reaction (toothpicks and gumdrops or clay work well if you do not have ball-and-stick models). Use the molecular level graphics of chemical reactions in Section 3.9 to aid your discussion.
- The text lists three steps for writing and balancing chemical equations. Step 3 tells the students to start "with the most complicated molecule(s)." For the reaction



consider balancing the equation in the following way:



Notice that we eventually get the right answer (so trial and error is a reasonable method). However, it would have been better not to have balanced the oxygen as the second step. This is why Step 3 tells the students to start with the most complicated molecule. Encourage the students to leave the elements and diatomic molecules until last when balancing equations. This simplifies the trial and error process.

- Emphasize that the coefficients in a balanced equation represent the mole *ratios*. An individual coefficient is meaningless (just as an amount of a single ingredient in a recipe is meaningless without the other amounts).
- It is useful to use a real life example when introducing limiting reactants. Students seem to relate easily to kitchen experiences yet have difficulty applying the same concepts to chemical equations. By working with real life experiences first and addressing the concepts in this context students have an opportunity to become comfortable with the concept before applying it to an unfamiliar setting.
- Using a concrete analogy (such as making sandwiches) helps the students better understand the calculations. This is especially true for problems involving a limiting reactant. Problems in which a reactant is limiting are conceptually the same as the cases in which no reactant is limiting. However, students seem to have much more difficulty dealing with limiting reactants. For example, some students calculate the amount of product that could be made from each reactant and then add them. Or, they add the moles of each given reactant for the total. They almost never make these types of mistakes in the analogies, though, so it is helpful to be very explicit about these points.
- Use molecular diagrams like those shown in Section 3.11 of the text to make counting molecules more concrete (and similar to real life experience).

CHAPTER FOUR: TYPES OF CHEMICAL REACTIONS AND SOLUTION STOICHIOMETRY

Chapter Learning Goals:

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|----------------|---|
| Section One: | To show why the polar nature of water makes it an effective solvent. |
| Section Two: | To characterize strong electrolytes, weak electrolytes, and non-electrolytes. |
| Section Three: | To define molarity and demonstrate calculations involving the composition of solutions. |
| Section Four: | To introduce several types of solution reactions. |

- Section Five: To show how to predict whether a solid will form in a solution reaction.
- Section Six: To describe reactions in solution by molecular, complete ionic, and net ionic equations.
- Section Seven: To demonstrate stoichiometric calculations involving precipitation reactions.
- Section Eight: To show how to perform calculations involved in acid-base volumetric analysis.
- Section Nine: To characterize oxidation-reduction reactions.
To describe how to assign oxidation states.
To identify oxidizing and reducing agents.
- Section Ten: To describe the oxidation states method for balancing oxidation reduction reactions.

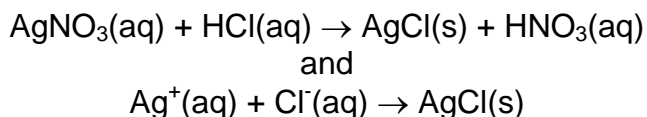
The placement of this chapter differs from its placement in many general chemistry texts. We feel there are several good reasons for including these topics at this point and many other texts have come to agree with us. In the eighth edition there is an emphasis on representing molecules and ions in solution.

A thorough discussion of the types of reactions in solution allows for an early introduction of descriptive chemistry. Descriptive chemistry is very important in the first-year course. Although virtually every instructor will give a different definition of descriptive chemistry and a different list of topics to be covered, a thorough discussion of chemical reactions in aqueous solution is probably central to all of those definitions and lists. Hence, the expanded discussion of reactions at an early point.

Several years ago texts began introducing molarity and solution stoichiometry in the stoichiometry chapter to give flexibility to the lab program. The expanded discussion in Chapter 4 gives a great deal of flexibility to the lab. The table on scheduling lectures and labs in Part I amply illustrates this point. The stoichiometry for all of the common types of solution reactions is discussed. Students can begin to see all types of reactions in the lab very early in the course. The expanded discussion in the text will help the student focus on what is actually going on in a solution, what species are really present, and how they interact with one another.

A chemical equation written on the page of a textbook probably has little meaning to a neophyte. Students must have the opportunity to see reactions. Lecture demonstrations and the color photographs in the text help, but the most useful place for seeing chemistry is by doing chemistry in the laboratory. Chapter 4 expands the vista of what can be done in the lab.

In Chapter 4, reactions are classified as precipitation, acid-base, and oxidation-reduction. The use of terms such as metathesis, combination, displacement, double displacement, and so forth, is avoided because these terms are not used by chemists. These terms can be confusing. For example, are



the same or different? Is the first a double displacement and the second a combination? We strongly feel reactions should be classified on the basis of what can be observed in the lab and what is happening chemically, and not on how an equation is written on paper.

We think that an early and thorough discussion of the topics in Chapter 4 will lay a strong foundation for integrating descriptive chemistry with chemical principles throughout the rest of the course. However, instructors may wish to postpone some of the topics until later. The sections on molarity and solutions in general can be used as part of a discussion of stoichiometry or postponed until Chapter 11 is covered. Acid-base and precipitation reactions can be delayed until Chapters 14, 15, and 16. The oxidation-reduction section can be covered with Chapter 18, Electrochemistry. The sections are quite independent and delay will not cause disruption of continuity.

At the beginning of the end-of-chapter exercises in Chapters 11 and 18, there are some problems that review the topics in Chapter 4. Thus, if you decide to do all of Chapter 4 early, the students can review the material later by doing these problems. We recommend doing the entire chapter early if at all possible. We think you will find it quite effective as a complement to the lab, which really must become a more important component of the general chemistry course.

Teaching Tips

- Molarity, like density (and molar mass), is a ratio of two numbers. Students often believe that if, for example, solution A has a greater concentration than solution B, then solution A must have a greater number of moles of solute than solution B. This is not necessarily the case. For example, present the following situation:

Solution A: dissolve 5 moles of sugar in 10 L of solution: concentration = 0.5 M

Solution B: dissolve 1 mole of sugar in 1 L of solution: concentration = 1.0 M

Solution B has fewer moles of sugar, but a greater concentration than solution A.

- The key concept in dilution is that the concentrated solution has the same number of moles as the dilute solution. Students are confused about this because they think of dilute as containing less solute than concentrated. In this case the solution has increased in volume and the number of moles of solute have remained the same. Stress that concentration is the amount of solute per volume available.

- The formula $M_1 V_1 = M_2 V_2$ works well but some students memorize it without understanding why it works. Emphasize what each of the variable means, and that it only works because the number of moles of solute is the same before and after dilution. Consider asking a question in which the students are asked to solve for the volume of water that must be added to achieve a dilution. In this way, the students need to think about what they are solving for instead of just plugging numbers into an equation.
- Students often conserve subscripts in precipitation reactions. Emphasizing what the solutions look like at a molecular level, and that the products must be electrically neutral, helps the students understand why subscripts need not be conserved.
- Discuss the importance of oxidation-reduction reactions in daily life. Batteries, corrosion, and combustion are all oxidation-reduction reactions.
- Point out to the students that the loss or gain of electrons can only take place if there is another element present to accept or donate electrons.
- Some students wonder why the term “reduction” is used to describe a *gain* of electrons. When you discuss oxidation states in Section 4.9, emphasize that when an atom is reduced its oxidation state is reduced.
- It is easiest to introduce the oxidation states for ions before discussing covalent molecules. Since the charge on the ion is the same as the oxidation state, this makes sense to the students. Emphasize that the charge on an ion is acquired from a transfer of electrons. This establishes the idea of assigning electrons to a particular atom. Then move to covalent molecules and begin assigning oxidation states.
- Students do need to learn the rules for assigning oxidation states. It is a good idea to give them a number of practice problems in assigning oxidation states to all of the elements in a compound.

CHAPTER FIVE: GASES

Chapter Learning Goals:

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| Section One: | To demonstrate atmospheric pressure and explain how barometers work.
To define the various units of pressure. |
| Section Two: | To describe certain laws that relate the volume, pressure, and temperature of a gas and to do calculations involving these laws. |
| Section Three: | To define the ideal gas law. |

- | | |
|----------------|---|
| Section Four: | To show how to do calculations involving the ideal gas law.
To define the molar volume for an ideal gas.
To define STP.
To show how to do stoichiometric calculations for reactions involving gases. |
| Section Five: | To show how to calculate molar mass from gas density.
To state the relationship between partial pressures and total pressure and between partial pressure and mole fraction.
To show how to obtain the molecular formula, given the empirical formula and the molar mass. |
| Section Six: | To present the basic postulates of the kinetic molecular theory.
To define temperature.
To show how to calculate and use root mean square velocity. |
| Section Seven: | To describe effusion and diffusion.
To show the relationship between effusion and diffusion. |
| Section Eight: | To describe how real gases deviate from ideal behavior.
To show how van der Waals's equation allows for real conditions. |
| Section Nine: | To characterize several real gases. |
| Section Ten: | To characterize the composition of the atmosphere.
To describe some of the chemistry of air pollution. |

The placement of this chapter is consistent with an historical approach. The studies of gases were fundamental in the development of the modern atomic theory. Boyle's law is consistent with an atomistic view of matter and not the Aristotelian view of matter. This material also provides an excellent example of how models are developed and used, an emphasis throughout the text. Consistent with the emphasis of other chapters, in the eighth edition we have increased the pictorial representations of gases and their reactions.

This chapter discusses the bulk of behavior of gases (gas laws and their applications) and follows with a microscopic model (the kinetic molecular theory) that describes this bulk behavior. The kinetic molecular theory is discussed in a more quantitative manner in Appendix Two. The behavior of real gases can lead logically to a discussion of atomic and molecular structure. One of the causes for deviation from ideal behavior is the interactions arising from interparticle attractions. To understand these attractions we must consider the detailed electronic structure of atoms and molecules rather than treat them as small, billiard ball-like particles.

The other logical place to deal with gases is to cover them between Chapters 9 and 10. After we discuss the microscopic structure of matter, we deal with the bulk physical properties of matters. Gases, liquids, and solids can be taken together at this point.

Teaching Tips

- This chapter emphasizes the scientific method in action. Focus on the way that observations lead to natural laws, and how we try to explain these laws with theories or models. This is a good time to review the difference between laws and models. The gas laws allow students to predict the behavior of gases, while the kinetic molecular theory explains why gases behave in the manner they do.
- It is important for the students to understand that while the laws are based on observations the laws do not explain why the observation occurs. Laws can help us make predictions, but they are not explanations. Knowing the law of gravity, for example, allows us to predict that if we drop a pen it will fall to the ground. It does not explain why the pen fell and does not explain to us what gravity is, or how gravity “works.” To explain the observations, we develop models or theories. We want the model to be as simple as it can be in order to answer the questions we wish to answer.
- You cannot overemphasize that for Boyle’s law to hold true the temperature and amount of gas must be held constant. Certain activities, such as filling a tire or a basketball with air, seem on the surface to be at odds with Boyle’s law. Students often reason this way: “If I add more gas to a container such as a basketball, the volume and pressure both increase. Thus, as pressure increases, so does volume.” Because of this, students may believe that pressure and volume are directly related.
- Students sometimes confuse the idea of an inverse relationship with that of a line having a negative slope. A linear graph with a negative slope is a direct relationship between the variables. To make this distinction clear for these students, focus on the idea that inverse means the relationship varies as the reciprocal ($1/x$). Have the students focus on the shape of the graph in Figure 5.5a. Ask the students to tell you what will happen to the pressure as the volume becomes extremely large (the pressure approaches zero, but will never reach zero). As the volume gets smaller, the pressure increases (as the volume goes to zero, the pressure goes to infinity). Thus, the graph must not touch the axes—it is a hyperbola not a straight line.
- At this point all four variables (P , V , T , and n) have been used in gas law problems. This is an excellent time to focus on problem solving skills. Be sure to have the students write down what they know from the problem, what they are looking for, and what they need to know to find it.
- Encourage the students to use the Ideal Gas Law to solve most gas law problems, even when some conditions are not changing. All the other gas laws can be derived from the Ideal Gas Law. If students understand how to do this they need not memorize all the other gas laws. Encourage the use of finesse and

understanding to replace the brute force method of memorizing formulas and plugging numbers into them.

- When introducing Dalton's Law of Partial Pressures, encourage students to think about the mixtures of gases at the molecular level. The presence of additional gases does not affect the number of times a particular gas molecule hits the side of the container (since it is mostly empty space anyway). This helps the students understand why the pressures are additive. The observations summarized by Dalton's Law of Partial Pressures are extremely important when developing a theory to explain the gas laws and we will develop a theory of gases that is dependent on these two statements.
- The postulates for the kinetic molecular theory form the basis for understanding the behavior of gases. Encourage students to develop a mental movie of an ideal gas on a molecular level that can form the basis for understanding the theory.
- It is also advisable to stress that we are talking about ideal gases. For example, we know real gas particles exert forces on one another as evidenced by steam condensing on a cool mirror during a hot shower. However, it is important for students to understand that although our model for gases assumes ideal behavior and gases are not ideal we can use the model to predict behavior and understand gases. The difficulties with real gases surface in small volumes and at large pressures. The ideal gas law still allows us to predict the behavior of gases under most everyday conditions.
- Once students see the molecular level action of gases with changing temperature proceed to investigate and develop models for the molecular level to illustrate the P, T and V, T relationships. This is an opportune time for the students to begin to think as a chemist. Observing gases at the macroscopic level allows prediction of behavior while developing a model at the microscopic level allows us to understand why gases behave as they do.
- Consider discussing a semi-quantitative relationship between pressure and volume for mathematically inclined students. Imagine a sample of gas in a cubic container with sides of 1 ft. If we transfer this sample to a cube with sides of 2 ft, the gas particles will have to travel (on average) twice as far from one wall to the next, and there will be half as many collisions (decreasing the pressure by a factor of 2). In addition, the area of each wall will be four times as large (4 ft^2 vs. 1 ft^2). Because pressure is a measure of force per area, this will cause the pressure to be decreased by a factor of 4. Thus, the pressure should be reduced overall by a factor of 8 (2×4). This is consistent with Boyle's law since the second container has a volume 8 times larger than the first container.

CHAPTER SIX: THERMOCHEMISTRY

Chapter Learning Goals:

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| Section One: | To describe the energy flow between a system and its surroundings.
To discuss the first law of thermodynamics.
To show how to calculate the work that results from changing the volume of a gas at constant pressure. |
| Section Two: | To define enthalpy and demonstrate calculations of the change in enthalpy in a chemical reaction.
To show how a change in enthalpy is measured by calorimetry. |
| Section Three: | To discuss the characteristics of enthalpy changes.
To show how to calculate ΔH for a chemical reaction. |
| Section Four: | To define standard states.
To show how to use standard enthalpies of formation to calculate ΔH° for a reaction. |
| Section Five: | To discuss fossil fuels and the effects of their use on climate. |
| Section Six: | To discuss energy alternatives.
To compare the available energy of various fuels. |

Once again, there are other logical alternatives to the placement of this chapter. Many instructors like to treat all of thermodynamics at once. Thus this chapter can be combined with Chapter 17 and covered either here or in the second semester of the one-year course.

We have tried to strike a compromise. Students usually have an intuitive feel for energy. Putting that intuition into a formal discussion is very helpful. In the lab the students will often encounter reactions involving the generation or absorption of heat. Energy is often involved in the processes discussed in connection with atoms and bonding (e.g., ionization energy, electron affinity, bond energy). In addition, a discussion of energy allows inclusion of some interesting and important descriptive chemistry, such as that given in the chapter on the energy crisis and alternative fuels. Some discussion of energy is useful in the first part of the year. We have included the first law (conservation of energy) and enthalpy in Chapter 6. We can then talk about energies, thermodynamic cycles (such as the Born-Haber cycle), and bond energies in Chapters 7 and 8. Topics that students find more difficult conceptually, such as entropy and free energy, are deferred until later in the text.

We recommend two things for instructors who choose to cover all of thermodynamics later in the course. First, look at heat as a stoichiometric quantity and present a general discussion of energy (Sections 6.1 and 6.2). Energy concepts are used heavily in Chapters 7, 8, 10, and 11; a brief introduction to energy would make these later discussions more clear. Also, consider saving the detailed discussion of bond energies (Section 8.8) for the later, comprehensive coverage of thermodynamics.

Teaching Tips

- Another example of a state function is displacement. For example, if two students go from Chicago to New York, they are both the same distance away from where they started initially. However, they may have taken different routes; one student may have traveled a greater distance, but they end up with the same displacement.
- Students have difficulty making a distinction between temperature and heat. It is important to talk about heat as a flow of energy and temperature as a measure of the “vigor” of the random motions of the components of a substance. Thermal energy is the energy a substance contains due to the random motions of its components. Heat is the way that thermal energy is transferred from one object to another.
- Thermodynamic quantities have two parts: the magnitude and the sign. It will be a new idea for many of the students to think of the sign as a representation of direction. This idea is a key concept on which to focus attention. Emphasize that the sign convention is arbitrary. That is, since chemists chose to take the system's point of view, all signs are determined relative to the system.
- Enthalpy has no physical meaning; it is a mathematic definition. At constant pressure, enthalpy shares the same meaning as heat.
- Once students understand that energy changes can be measured in a calorimeter, the next logical question is: “How can anyone measure all of the energy changes in the world? What if the reaction is unpleasant or results in an explosion?” To answer this question, Hess’s law is useful. It provides an excellent example of how a chemist can extend knowledge by applying information gathered from simple systems to more complex systems.
- An analogy you can use with Hess's law is paying for something at a store. If the net result is to pay the store \$8, this can be achieved in many ways. For example, you can give the store one \$5 bill and three \$1 bills; or you can pay with a \$10 and get \$2 in change. No matter how many steps are involved in the process, the payment is the same.

CHAPTER SEVEN: ATOMIC STRUCTURE AND PERIODICITY**Chapter Learning Goals:**

Section One: To characterize electromagnetic radiation in terms of wavelength, frequency, and speed.

Section Two:	To introduce the concept of quantized energy. To show that light has both wave and particulate properties. To describe how diffraction experiments were used to demonstrate the dual nature of all matter.
Section Three:	To show that the line spectrum of hydrogen demonstrates the quantized nature of the energy of its electron.
Section Four:	To describe the development of the Bohr model for the hydrogen atom.
Section Five:	To show how standing waves can be used to describe electrons in atoms. To describe the Heisenberg uncertainty principle. To explain the significance of electron probability distributions.
Section Six:	To explain the quantum numbers n , l , and m_l .
Section Seven:	To describe the shapes of orbitals designated by s, p, d, and f and to discuss orbital energies.
Section Eight:	To define electron spin and the electron spin quantum number. To explain the Pauli exclusion principle.
Section Nine:	To show how the quantum mechanical model can be applied to atoms besides hydrogen.
Section Ten:	To trace the development of the periodic table.
Section Eleven:	To explain the Aufbau principle.
Section Twelve:	To show general trends in ionization energy, electron affinity, and atomic radius in the periodic table.
Section Thirteen:	To show what types of information can be obtained from the periodic table.

This chapter combines into one what many texts do in two. Particular attention is paid to the relationship between models and facts. The order of topics is very similar to texts that treat electronic structure and the periodic table separately. The nature of light and the quantum mechanical model of the hydrogen atom are discussed. Then the historical development of the periodic table is followed by a discussion of the quantum mechanical model of polyelectronic atoms and how the model fits the periodic table. In discussing polyelectronic atoms, observation and the periodic table are emphasized. The concept of shielding is used to rationalize the periodic trends in size, ionization energy, and electron affinity. At first this may seem to be too sophisticated and abstract an approach for first-year chemistry students. However, the approach is very pictorial and easy to comprehend. We have found this approach to work well in our classes. Emphasis can be shifted at the instructor's option between models and the experimental facts. Combining the atomic theory with the periodic table makes this easier.

The last part of the chapter discusses the periodic trends and properties of the alkali metals. This presents the instructor with another opportunity to integrate more descriptive chemistry in with the chemical principles. Chapter 20 is organized to discuss the elements by periodic group. Instructors may choose other groups from these chapters to illustrate chemical periodicity.

A further point should be made concerning atomic radii. The volumes given in Chapter 7 are for covalent radii. The covalent radii of He, Ne, and Ar are estimates. All comparisons of atomic sizes to each other can be done consistently with this set of radii.

In Chapter 8, instructors may want to compare the sizes of a series of ions to the isoelectronic noble gases. The radii of the noble gases for this comparison should be the univalent crystal radii. These values are: He = 93 pm, Ne = 112 pm, Ar = 154 pm, Kr = 169 pm, and Xe = 190 pm. Huheey¹ and Pauling² provide detailed information on uses of ionic radii and tables of their values.

Teaching Tips

- Students often consider the forms of electromagnetic radiation outside the visible range (radio, X-rays, infrared, microwaves) to be entirely different from light. Use Figure 7.2 (the electromagnetic spectrum) to show the range of frequencies for each type of radiation.
- A commonly used analogy for the energy levels within an atom is the staircase. A person can move from one step to another or even move up 2 or 3 steps and down a similar number. It is not possible however to move up or down a part of a step. Within atoms electron energies are quantized—the energy levels are like stairsteps. Electrons can change only between established energy levels, not in between. Compare this to using a ramp which is like continuous energies.
- Be sure to discuss how hydrogen can have multiple lines even though it has only one electron. Students often forget that they are observing multiple atoms simultaneously when observing a line emission spectrum.
- Point out that the quantized nature of atoms is a surprise and non-intuitive. It is comforting to students to realize that this topic is complicated even to people who have been studying it for awhile.
- This is an excellent time to discuss the colors of streetlights. The mercury vapor lights have a blue cast, while the sodium lights are yellow and neon lights are more orange. You may also want to discuss the difference between an incandescent light and a vapor light. Many students think all lights are the same. They may not realize that an incandescent lamp produces a continuous spectrum since the tiny filament is radiating all wavelengths of visible light while vapor

¹James E. Huheey, *Inorganic Chemistry; Principles of Structure and Reactivity*, 3rd ed., Harper and Row (1983).

²Linus Pauling, *The Nature of The Chemical Bond*, Cornell University Press (1960), pp. 511-519.

lamps (like many street lights) are more like gas discharge tubes in their operation.

- It is a good idea to emphasize that although the Bohr model of the atoms is important historically, it does not represent an accurate model for the atom. Bohr's model is no longer accepted because it is considered fundamentally incorrect (this is why the section is so short). A discussion of this model is useful in discussing the nature of science. The Bohr model is a relatively simple model with the sole intent of explaining the results of the hydrogen emission spectrum. It succeeded in doing so and this is why it was considered powerful. However, this model does not fit polyelectronic atoms and electrons do not move in fixed orbits.
- In this section the idea of probability as a model for understanding nature is introduced. This will be new to all students and most will find it confusing. It is a fundamental change from determinism in looking at the world and it is understandable that students will have difficulty with it. For example, students will tend to think of orbitals as physical structures and will find it difficult to think of them as probability distributions. In Section 7.5 we describe the orbital as a *potential space* for an electron. This is a good way for the students to think about it initially.
- Many students confuse orbitals and orbits. An orbit describes a particular path an object follows as it travels around another. For example, the moon has an orbit about the earth. Electrons do not follow a particular path around the nucleus. An orbital describes the volume around the nucleus where an electron is likely to be found. The exact path of an electron in this area is not known.
- For polyelectronic atoms the third electron goes to the 2s orbital. In the hydrogen atom the 2s and 2p orbitals have the same energy. However, multiple electrons change this. With multiple electrons, in a given energy level the s orbitals are lower in energy (and thus filled first) than the p orbitals, which in turn are lower in energy than the d orbitals.
- Encourage the students to use the periodic table to assist them in figuring out electron configurations. Once they see that they can just move across the rows, filling electrons into the s and p orbitals, electron configurations become much easier to write. This should be used to emphasize the periodicity of the periodic table (even though the table was invented before the discovery of electrons). Thus, the students should realize that with an understanding of how the periodic table is put together, the students can figure out the expected electron configuration of any element. They need not memorize these. See Figure 7.28 in the text.

- Make sure students understand the trends for atomic size across rows and down groups in the periodic table. If they understand these trends, the trends for ionization energy are much easier to grasp. The students should be able to see that the trends for atomic size and ionization energy are consistent with one another.
- Many students believe that alkali metals “want” to lose an electron; that is, the potassium atom, for example, will release energy to become an ion. Or they believe that less energy is required to remove the second electron from calcium, for example, than the first because calcium “wants” to lose two electrons. Energy is always required to remove an electron, and successive electrons can only be removed with an increasing amount of energy.

CHAPTER EIGHT: BONDING: GENERAL CONCEPTS

Chapter Learning Goals:

Section One:	To explain why an ionic bond is formed. To explain why a covalent bond is formed. To introduce the polar covalent bond.
Section Two:	To discuss the nature of a bond in terms of electronegativity.
Section Three:	To define the relationship between bond polarity and molecular polarity.
Section Four:	To show how to predict the formulas of ionic compounds. To discuss the factors governing ion size.
Section Five:	To define lattice energy and show how it can be calculated.
Section Six:	To show the relationship between electronegativity and the ionic character of a bond.
Section Seven:	To discuss the covalent bonding model.
Section Eight:	To show how bond energies can be used to calculate heats of reaction.
Section Nine:	To introduce the localized electron model.
Section Ten:	To show how to write Lewis structures.
Section Eleven:	To show how to write Lewis structures for certain special cases.
Section Twelve:	To illustrate the concept of resonance. To show how to write resonance structures.
Section Thirteen:	To describe how molecular geometry can be predicted from the number of electron pairs.

Chapter 8 is the first of two chapters that covers chemical bonding. This arrangement is typical of most current general chemistry texts, but there are some important differences. One difference is the approach used in Chapter 8 and the second is the inclusion of VSEPR in this first chapter on bonding, rather than its placement in the

chapter on orbitals.

The approach of Chapter 8 is to begin with a discussion of measurements that can be made that tell us something about the characteristics of bonds. Electronegativity is introduced so that measured dipole moments can be correlated to bond type. Ionic bonds are then discussed. The discussion of covalent bonding follows. The first section on covalent bonding begins with bond energies. Bond energies are, again, a quantity that can be measured. An emphasis is placed on the priority of measurement over theory, something that is lost in many books.

The use of VSEPR theory to predict molecular geometry is included in Chapter 8. We feel this is more logical for two reasons. First, molecular geometry is discussed in the same chapter as polarity. Second, this placement gives the instructor greater flexibility.

A key concept a student should get from the first-year course is to appreciate the intimate relationship between molecular structure and the bulk physical and chemical properties of matter. Including VSEPR in Chapter 8 places all of the topics necessary for dealing with this correlation in a single chapter. Instructors can then go directly to Chapter 10 and emphasize the structure-property relationships, leaving the more abstract topics on orbitals in bonding for a later time.

The section on formal charge should be considered optional. We feel that a discussion of formal charge is not essential in a first-year course and that there is a danger in overusing the concept. The section and end-of-chapter exercises are organized so that formal charge can be easily skipped. If you do choose to cover formal charge, we urge caution and suggest that you try to limit the discussion to using formal charge for deciding on reasonable resonance structures.

Teaching Tips

- Ionic and covalent bonding are the extremes of types of bonding. It is important that students understand that attractive forces between molecules lie on a continuum.
- Focus the student's attention on where the electrons "spend more time". Are they equally shared between two atoms or do they spend more time closer to one atom in the bond? This focus will help the students understand bond polarity.
- The relationship between the difference in electronegativity and bond type is rather arbitrary and only serves as a rough guide. Thus, the transition between polar covalent and ionic bonding, for example, is not a sharp one and the students should know this.
- Students will often confuse dipole moments in a molecule with the overall polarity of a molecule. Molecules with polar bonds may be non-polar molecules (CCl_4 is

an example). The three-dimensional shape affects the overall polarity of a molecule (this is discussed in more detail in Section 8.13)

- The polarity of a molecule has a profound impact on its chemical and physical properties. For example, water is a liquid and methane (CH_4) is a gas at room conditions even though both have similar molar masses. Relating molecular shape to polarity to properties provides a “big picture” so the students know where we are going with this information.
- This is an excellent time to discuss the importance of polarity in molecules by using the properties of water. Discuss the ability of water to dissolve polar and ionic substances and the strong attraction that water molecules have for each other which results in a high boiling point for water.
- This is an excellent time to review the electron configurations for atoms and write electron configurations for ions. Emphasize that we can predict the ion that forms from a particular atom by achieving a noble gas electron configuration. The concept that there is a strong tendency for atoms to adopt a noble gas configuration is central to understanding the electronic structure of ions and the Lewis structures of molecules. This concept is emphasized in this chapter and should be stressed in class discussions. However, this does not explain *why* the atom achieves this configuration. Science is about observing patterns (laws) and developing explanations (theories or models). The fact that the atoms in a compound have a tendency to achieve a noble gas electron configuration is an observation, not an explanation. Knowing this helps us to make predictions, but it does not answer the “why?” question. Knowing this observation, students should be able to predict, for example, that a sodium atom will lose only one electron when becoming an ion, and that a calcium atom will lose two electrons. However, the reason is not because the atoms “want” to have a noble gas electron configuration. This is a subtle but important distinction.

At the beginning of this discussion it should be emphasized that when chemists use the term *ionic compound*, they are usually referring to the solid state of that compound. In the solid state the ions are close together. That is, solid ionic compounds contain a large collection of positive and negative ions packed together in a way that minimizes the $- \dots -$ and $+ \dots +$ repulsions and maximizes the $+ \dots -$ attractions. This situation stands in contrast to the gas phase of an ionic substance, where the ions are quite far apart on average. In the gas phase, a pair of ions may get close enough to interact, but large collections of ions do not exist. Thus, when we speak of the stability of an ionic compound, we are referring to the solid state, where the large attractive forces present among oppositely charged ions tend to stabilize (favor the formation of) the ions. For example, the O^{2-} ion is not stable as an isolated, gas-phase species but, of course, is very stable in many solid ionic compounds. That is, MgO(s) , which contains Mg^{2+} and O^{2-} ions, is very stable, but the isolated, gas-phase ion pair

$\text{Mg}^{2+} \quad \text{O}^{2-}$ is not energetically favorable in comparison with the separate neutral gaseous atoms. Thus you should keep in mind that in this section, and in most other cases where we are describing the nature of ionic compounds, the discussion usually refers to the solid state, where many ions are simultaneously interacting.

- You can effectively show why noble gases do not generally form compounds by drawing the Lewis structures for several noble gases. However, it is a good idea to de-emphasize the drawing of Lewis structures for atoms. It is shown here to introduce simple Lewis structures (diatomic molecules in particular). But students are later confused when confronted with a molecule such as carbon monoxide (CO), for which we cannot draw the Lewis structure of the molecule directly from the Lewis structures of the atoms.
- Be sure to emphasize that in order to write correct structures for molecules it is necessary to sum all the valence electrons from all the atoms. Be sure that students do not attach valence electrons to individual atoms and then try to write a Lewis structure. This method can sometimes produce a correct structure but often does not for molecules containing multiple bonds and for complex molecules.
- Remind students that many elements do usually obey the octet rule. If an atom that can exceed the octet rule is bonded to atoms that obey the octet rule, the available electrons should be distributed to form octets on all the atoms. Then any remaining electrons can be placed on atoms that can exceed the octet rule.
- Use models in class to demonstrate each of the shapes in this section. It is difficult for students to visualize the shapes of molecules without models. Students can make models at home to help them with problems by using toothpicks and gumdrops or marshmallows. There is an activity in which the students make these shapes with balloons in the *Inquiry Based Learning Guide*.
- When teaching the VSEPR model, be sure to stress the difference between the arrangement of the electron pairs and the molecular structure. For example, point out the differences between BeCl_2 and water. This will help the students see the difference between electron pair structure and molecular shape. Many students do not understand why water should have a bent shape and BeCl_2 is linear.
- It is crucial that the students write the Lewis structure first. Many students believe they can predict the shape from the formula. However, it is easy to show counterexamples of this. For example, CO_2 is linear, while OF_2 is bent; BF_3 is trigonal planar, while NH_3 is trigonal pyramid.

CHAPTER NINE: COVALENT BONDING: ORBITALS

Chapter Learning Goals:

- | | |
|----------------|---|
| Section One: | To show how special atomic orbitals are formed in covalent bonding. |
| Section Two: | To show how molecular orbitals are formed in a molecule.
To define bond order and demonstrate how to calculate it. |
| Section Three: | To discuss the bonding in certain molecules of the general formula X_2 .
To relate paramagnetism to the filling of molecular orbitals.
To correlate bond order, bond energy, and bond length. |
| Section Four: | To use the molecular orbital model to treat bonding between two different atoms. |
| Section Five: | To show how the need for resonance is eliminated if the localized electron and molecular orbital models are combined. |

With the inclusion of geometry and VSEPR in Chapter 8, this chapter deals strictly with orbitals and bonding. Both valence bond and molecular orbital models are discussed. Instructors who wish, either because of time constraints or philosophy, to do so can delay this chapter to later in the second semester or use parts as appropriate.

The end-of-chapter exercises try to emphasize some of the more practical applications of the bonding models. For example, consideration of the nature of the π -bonds leads one to the conclusion that C_2H_4 is planar and C_3H_4 is not. The exercises on M.O. theory also emphasize the relationship between the model (predicted bond orders) and observations (bond energies, bond lengths, and paramagnetism).

An additional topic instructors may wish to introduce at this point is the difference between second period elements and their heavier congeners. Much of the difference between Si and C (Chapter 20), for example, can be attributed to differences in bonding.

Teaching Tips

- The idea of hybridization and molecular orbitals highlights the concept that molecules are not simply “atoms stuck together.” The atomic orbitals change to accommodate the formation of molecules.
- A discussion of hybridization is another area in which you can discuss the nature of science. In this case, we create a model based on empirical evidence (observations). The fact that the bond angles in methane, for example, are not 90° leads us to think about how atomic orbitals must change in the formation of compounds.

- Figures 9.5, 9.9, and 9.16 show the students to use “lowest energy arguments” when determining the stability of a molecule. This complements the discussion that began in Chapter 6.
- Point out Figure 9.24 to students. This is a good summary to show the students that hybridization is dependent on geometry, which comes from the Lewis structures. So, if you ask a student to determine the hybridization of a certain atom in a molecule, they are not expected to just “know” it—they must write out the Lewis structure first.
- Some students will find it very helpful to make physical models of figures such as Figure 9.13 and Figure 9.19.
- The molecular orbital model relies on the fact that orbitals are wave functions and that waves can constructively interfere or destructively interfere. Thus, we get bonding or antibonding orbitals, respectively.
- Use Figure 9.38 to show how bond order (from MO theory) is related to the number of bonds in a Lewis structure. This shows that the models are consistent.
- Figure 9.26 (and others like it) shows not only the molecular orbitals, but the atomic orbitals to show relative energies. Figure 9.38 shows only the molecular orbitals.

CHAPTER TEN: LIQUIDS AND SOLIDS

Chapter Learning Goals:

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|----------------|---|
| Section One: | To define dipole-dipole force, hydrogen bonding forces, and London dispersion forces.
To describe the effects these forces have on the properties of liquids and solids. |
| Section Two: | To describe some properties of liquids: surface tension, capillary action, and viscosity. |
| Section Three: | To contrast crystalline and amorphous solids.
To introduce X-ray diffraction as a means for structure determination. |
| Section Four: | To discuss the concept of closest packing of metal atoms.
To describe two models for bonding in metals.
To define and classify alloys. |
| Section Five: | To show how the bonding in elemental carbon and silicon accounts for the widely different properties of their compounds.
To explain how a semiconductor works. |

- Section Six: To describe the bonding in molecular solids.
Section Seven: To model the structures of ionic solids using the packing of spheres.
Section Eight: To define the vapor pressure of a liquid.
To discuss the features of heating curves.
Section Nine: To discuss the features of phase diagrams.

Chapter 10 is one of the key chapters in the text. The correlation of structure and properties is one of the thought processes central to all of chemistry. Solid-state chemistry is becoming increasingly important.

One of the key thought processes used by chemists and one that is perhaps unique to chemists, the correlation between structure and properties, is the idea that permeates this chapter. From the nature of the bonding and the geometry of molecules, we can predict the nature and relative strengths of interparticle forces. With that in hand we can begin to either predict or correlate physical properties to structure. The same type of reasoning can be extended to rationalizing and categorizing chemical properties.

Topics such as types of crystals, unit cells, closest packing, vapor pressures, and phase diagrams can be given variable emphasis depending on the students' and instructor's interests. With a class of mostly engineers, for example, extra emphasis might be given to the solid state and particularly semiconductors. Or a discussion of metallurgy (Chapter 21) can be covered here.

Steelmaking and polymer additives (Chapter 21 and 22) are excellent examples of how the properties of materials are chemically fine-tuned. If the extensive section on polymer additives is discussed at this point, at least a brief discussion of organic compounds and nomenclature (Chapter 22) would be in order. Alternatively, the section on the differences between carbon and silicon can be expanded by dealing more with their compounds here (Chapters 20 and 22).

For example, at the University of Illinois at Urbana-Champaign, one of the second-semester general chemistry courses (Chemistry 104) is aimed primarily at engineering students. We begin the second semester by covering Section 10.1 on intermolecular forces and then spend about three weeks on organic chemistry and synthetic polymers, emphasizing structure-property relationships. Later we come back to the rest of Chapter 10 when we talk about condensed phases, placing an emphasis on the solid state and electrical properties of solids. This chapter is really a pivotal chapter in the book. The overall flavor of a course is greatly affected by how instructors choose applications to illustrate the topics in this chapter.

Teaching Tips

- Students may need to be reminded that energy is required to change phase even though the temperature is not changing during that time. The energy applied to the system is used to overcome the intermolecular forces.

- Remind the students that a phase change occurs when no temperature change occurs and no change of phase occurs when the temperature is changing.
- Be sure to help the students make a clear distinction between intramolecular forces (covalent and ionic bonds) and intermolecular forces.
- This is an excellent opportunity to review the main concepts from bonding and prepare the students to look at the forces that exist between molecules. It is especially helpful to review the concept of a dipole and what it means before moving into intermolecular forces. This is the time to tie in electronegativity and arrangement of atoms to polarity of molecules.
- Some students think that dipole-dipole interactions and hydrogen bonds are completely unrelated. Be sure to stress that all of these interactions result from polarity.
- Students are often confused by the term hydrogen bond. They have the misconception that the hydrogen bond is the O-H, N-H, or F-H bond and not the attractive force between the dipoles. We have used dashed lines to represent hydrogen bonds to help the students see this more clearly. Stress that a hydrogen bond is an attractive force between molecules.
- Students have trouble deciding when molecules should have hydrogen bonds between them. Use examples such as NH_3 , CH_3OH , and HF to show that in each case the hydrogen is bonded directly to the N, O, or F. Molecules such as $(\text{CH}_3)_3\text{N}$ do not form hydrogen bonds even though they contain both N and H because there is no hydrogen bonded to the N to produce the dipole needed for a hydrogen bond.
- It is hard for student to visualize what happens within a molecule to produce London dispersion forces. This is difficult to represent with a picture. It is easiest to begin talking about these forces by using small nonpolar molecules.
- London dispersion forces exist not only among noble gas atoms and nonpolar molecules, but also between all molecules. However, they are generally only important among noble gas atoms and nonpolar molecules because they are relatively weak.
- The idea that evaporation is an endothermic process is hard for many students to grasp. Be sure to take time to explain that the temperature of the liquid drops because the higher kinetic energy molecules are going to the gas phase. Energy flows in if the container is not insulated and the temperature of the liquid will then remain constant.

- The familiar scents of finger nail polish remover and rubbing alcohol tell us that these liquids evaporate relatively easily. They have relatively high vapor pressures.
- Vapor pressure is really the beginning of introducing the concept of equilibrium to students. The idea that molecules are changing between the gas and liquid phase at constant equal rates is new for students. Focus the attention of the students on Figure 10.38 to help them understand what is taking place at the molecular level.
- Remind students that the formulas for ionic solids are actually empirical formulas, not molecular formulas.

CHAPTER ELEVEN: PROPERTIES OF SOLUTIONS

Chapter Learning Goals:

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|----------------|--|
| Section One: | To define various ways of describing solution composition. |
| Section Two: | To define the heat of solution and discuss its various energy components. |
| Section Three: | To show how molecular structure, pressure, and temperature affect solubility. |
| Section Four: | To show how a solution's vapor pressure is affected by the concentration of solute and the interactions of solute and solvent. |
| Section Five: | To explain the effect of a solute on the boiling and freezing points of a solvent. |
| Section Six: | To explain osmosis and describe its application. |
| Section Seven: | To show how the colligative properties of electrolyte solutions can be used to characterize the solute. |
| Section Eight: | To define a colloid and explain how it is stabilized. |

The treatment of solutions is fairly standard. An emphasis can be placed on structure-property correlation here, too. Enthalpy changes and thermochemical cycles can be discussed further by dealing with the energetics of solution formation. The remaining topics can be given whatever emphasis the instructor chooses.

Instructors who do not choose to do all of Chapter 4 early can begin introducing some of these topics (composition, nature of water, precipitation reactions) at this point in the course. It would also work well to go in the following order: Sections 4.1 to 4.3, Chapter 11, Sections 4.4, and 4.12. If Chapter 4 was covered in entirety early in the course, students can use Problems 1-8 in the end-of-chapter exercises as a review.

Teaching Tips

- A common example of a solution with a concentration expressed in mass percent is vinegar. You may want to bring a bottle to class showing a label giving the mass percent (usually 5%). You can have the students convert this to molarity, given the density.
- This is an excellent opportunity to repeat the conductivity demonstration used in earlier chapters. This will remind students that we can tell if a solution contains ions by its conductivity.
- Students can remember “like dissolves like” but often have trouble figuring out which substances are alike and which are not. Use the structures of similar molecules such as CH_3OH and water to show their similarities.
- Explaining why two non-polar substances mix is difficult without bringing in the concept of entropy. The text addresses this with a discussion of probability of mixing (see page 503).
- Although the text uses the term *hydrophobic*, the term is actually a misnomer. While a substance such as oil does not mix with water, a nonpolar molecule does not repel a polar molecule. In fact, they have attraction for one another (London dispersion forces). However, the attraction for a polar molecule with another polar molecule is greater, and it looks as though oil and water, for example, repel one another when we mix them together.
- Henry’s law and the discussion of vapor pressure introduce students to the idea of equilibrium (as did Section 10.8). This can be recalled when discussing chemical equilibrium in Chapter 13.
- The concentration unit molality is used for colligative properties because it is not dependent on a change in temperature (like molarity is).
- Note that the osmotic pressure formula, $\pi = CRT$ is analogous to the ideal gas law formula $P = (n/V)RT$. It turns out this stems from the model that vapor pressures are important when considering why a solution exhibits osmotic pressure.

CHAPTER TWELVE: CHEMICAL KINETICS**Chapter Learning Goals:**

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|----------------|---|
| Section One: | To define the reaction rate and to show how rates can be measured from experimental data. |
| Section Two: | To describe the two types of rate laws. |
| Section Three: | To learn methods for determining the rate law for a reaction. |

- Section Four: To develop rate laws relating concentration to reaction time and to show how they can be used to determine reaction order.
- Section Five: To explore the relationship between the reaction pathway and the rate law.
- Section Six: To discuss the temperature dependence of reaction rates.
To describe the collision model.
To define and show how to calculate activation energy.
- Section Seven: To explain how a catalyst speeds up a reaction.
To discuss heterogeneous and homogeneous catalysis.

Chemical kinetics seems to be one of the most difficult topics for students. Some of the topics, such as collision theory, are abstract—one of the reasons we have included an increased emphasis on pictorial representations of reactions in this and other chapters. Other topics, such as rate laws and graphing, require a good math background, a great deficiency in many beginning students.

Although this chapter is written to make this material understandable to students at this point in the course, some instructors may wish to postpone this chapter until after all of the equilibrium chapters are covered. Chapters 12 and 13 can also be reversed in order, as a full treatment of kinetics is not required before covering Chapter 13. Equilibrium is discussed from a phenomenological standpoint rather than on a kinetic basis. Another alternative order is to do kinetics with the exceptions of mechanisms first, then take up equilibrium, and finish with mechanisms. This approach allows one to cover in greater detail mechanisms that involve an equilibrium step. Instructors who covered nuclear reactions (Chapter 19) early with Chapter 2 can use the kinetics of nuclear decay as an example of a first-order reaction. Some of the end-of-chapter exercises cover nuclear decay. This allows for the mention of nuclear decay by instructors who do not cover Chapter 19 in detail.

Because we have placed kinetics before equilibrium, we have eliminated all end-of-chapter exercises that concern mechanisms with a fast equilibrium before the rate-determining step. Many instructors may wish to switch the order of these chapters.

Teaching Tips

- The study of chemical kinetics allows us to think about *how* chemical reactions occur. Stoichiometry tells us what happens (and amounts) and thermodynamics helps us to predict what will happen.
- Deriving the integrated rate law equations requires calculus.
- Watch for the misconception that a reaction is complete after two half-lives (in fact, the reaction is 75% complete after the first two half-lives). Use Figure 12.X when discussing this.

- Table 12.6 nicely summarizes the discussions of differential rate laws, integrated rate laws, and half-lives.
- Use Figure 12.10 to illustrate the energy changes that occur when 2 BrNO molecules from Figure 12.12 interact.
- Most collisions do not result in a chemical reaction. This is due to energy requirements and positioning of the molecules. If all collisions resulted in a chemical reaction, life would be extremely unstable. Virtually everything would happen at a much higher speed (including our aging). Students should appreciate the vast number and types of chemical reactions that are occurring all around them (especially in an atmosphere rich in oxygen).
- Mechanisms cannot be proven, but the observed rate laws can be used to support them.
- Remind students that the kinetic energy of a collection of molecules is directly proportional to the temperature (K).
- To get to Equation 12.11 in the text we must assume that A is not dependent on temperature. However, the value of A must be dependent on temperature since part of it is due to the number of molecular collisions. However, for relatively small temperature differences, this equation works (this is true because the number of collisions is relative to the square root of the temperature, and since must use a Kelvin scale, small differences in temperatures do not appreciably change the value of A).
- This is an excellent time to discuss the use of catalysts. Catalytic converters in automobiles use metal such as palladium and platinum to convert uncombusted hydrocarbons to carbon dioxide and water. The metals are used repeatedly in these processes. In the human body enzymes act as catalysts speeding up reactions that would take hours in the laboratory to fractions of a second in the body.

CHAPTER THIRTEEN: CHEMICAL EQUILIBRIUM

Chapter Learning Goals:

- | | |
|----------------|---|
| Section One: | To discuss how equilibrium is established. |
| Section Two: | To introduce the law of mass action and to show how to calculate values for the equilibrium constant. |
| Section Three: | To show how K and K_p are related. |
| Section Four: | To show how condensed phases are treated in constructing the equilibrium expression. |

- Section Five: To show how the equilibrium constant is used to predict the direction in which a system will move to reach equilibrium. To demonstrate the calculation of equilibrium concentrations given initial concentrations.
- Section Six: To generalize the procedure for doing equilibrium calculations.
- Section Seven: To show how to predict the changes that occur when a system at equilibrium is disturbed.

Students generally find the topic of chemical equilibrium to be quite difficult. One difficulty is that we treat reactions as macroscopically static (the concentrations do not change over time once equilibrium is reached) yet microscopically dynamic. Students do not always differentiate between the two, and usually visualize the reactions as static (the macroscopic is given precedence over the microscopic). This is why it is important to emphasize representations of the reactions, which is done in the text and asked of the students in some of the Active Learning Questions.

The chapter on chemical equilibrium is written to be a stand-alone chapter. It does not require that kinetics be covered first. There are exercises at the end of the chapter that cover the interplay between equilibrium and kinetics, and these can be assigned after both topics have been covered. The chapter introduces, primarily through gas phase examples, the phenomenon of equilibrium and the techniques that are useful in solving equilibrium problems. Appendix 1.4, Solving Quadratic Equations, can also be covered here. It describes the method of successive approximations that can be very useful in solving the equilibrium problems at the end of the chapter. Other examples of gas-phase reactions from other chapters can be brought in here to cover more descriptive topics. The synthesis of ammonia and the chemistry of nitrogen oxides (Chapter 20) or the petrochemical chemistry (Chapter 22) can be used as illustrations of gas-phase reactions.

Teaching Tips

- Point out that equilibrium can only be reached in closed systems. This includes gases in closed containers and ions in solution (since the ions cannot escape). In industry processes do not usually reach equilibrium. Instead the process is designed to maximize the amount of product formed (and avoid the limitation of equilibrium) by removing the product as it is formed.
- Emphasize that equilibrium is not a condition in which the reactants and products are present in equal amounts. The balance is between the rates of the forward and reverse reactions.
- Students often think that equilibrium is a static condition. Be sure to emphasize that, although at the macroscopic level things appear to be constant at equilibrium, at the molecular level chemical reactions continue to take place. The rates of the forward and reverse reactions are equal.

- Be sure to use balanced chemical equations including states for each chemical species in this section. In the next section we discuss heterogeneous equilibria and the students will need to distinguish among the states of the chemicals.
- Once students begin to understand how to setup the equilibrium expression for a reaction we can begin to talk about calculations involving the expression. It is important to distinguish between initial conditions and equilibrium conditions when students begin calculations.
- Have the students calculate the change in concentrations for each of the species in Table 13.1. Make sure the students understand that the ratio of concentration changes is the same ratio as the coefficients in the balanced equation.
- Use the results in Table 13.1 to emphasize the difference between “equilibrium position” and “equilibrium constant.”
- When calculating K from concentrations of reactants and products, it is important to emphasize both initial and equilibrium concentrations so that the students can distinguish between the two sets of data. This can be a major source of confusion for beginning students when working these problems.
- Students have trouble understanding how a change in volume affects the system. The idea that decreasing the volume increases the pressure and therefore changes the concentration is difficult for many students to understand.
- When working with gases it is important for the students to remember that they must use a balanced equation to make their predictions.
- For students who would like a mathematical verification of the effect of a change in volume, discuss how the partial pressures of the gases are changed. For example, suppose that at some temperature the equilibrium concentrations of N_2 , H_2 , and NH_3 are all 1.0 M (thus, at this temperature $K = 1.0$). If we suddenly double the volume, the concentrations of each species initially decrease to 0.50 M . If we put these numbers into the equilibrium expression we get a value of 4.0 . This number is higher than K so the system must shift to the left to satisfy the equilibrium condition ($K = 1.0$). This is the direction predicted by LeChatelier's principle.

CHAPTER FOURTEEN: ACIDS AND BASES

Chapter Learning Goals:

- Section One: To discuss two models of acids and bases and to relate equilibrium concepts to acid dissociation.
- Section Two: To relate acid strength to the position of the dissociation equilibrium.
To discuss the autoionization of water.
- Section Three: To define pH, pOH, and pK and to introduce general methods for solving acid-base problems.
- Section Four: To demonstrate the systematic treatment of solutions of strong acids.
- Section Five: To demonstrate the systematic treatment of solutions of weak acids.
To show how to calculate percent dissociation.
- Section Six: To introduce equilibria involving strong and weak bases.
To show how to calculate pH for basic solutions.
- Section Seven: To describe the dissociation equilibria of acids with more than one acidic proton.
- Section Eight: To explain why certain salts give acidic or basic solutions and to show how to calculate the pH of these solutions.
- Section Nine: To show how bond strength and polarity affect acid-base properties.
- Section Ten: To show how to predict whether an oxide will produce an acidic or basic solution.
- Section Eleven: To define acids and bases in terms of electron pairs.
- Section Twelve: To summarize the major species approach to solving acid-base problems.

This chapter emphasizes a problem-solving approach to acid-base chemistry first used in *General Chemistry Problem Solving II*, by Steven S. Zumdahl (D. C. Heath, 1979).³ The key features of the approach are to recognize which acids and bases are present, what reactions are possible, and how to break complicated problems down into more manageable ones.

In many instances the general chemistry course is the only exposure to aqueous equilibria that students will get. As a result, it is a topic that is covered fairly rigorously. The combination of chemistry with mathematical rigor gives many students a great deal of difficulty. Developing the students' ability to think things through in a systematic fashion will get them through acid-base equilibria with lessened anxieties and frustrations.

³These problem-solving techniques are also covered in *Solving Equilibrium Problems with Applications to Qualitative Analysis*, D. C. Heath, 1988. The chapters on equilibrium are the same in both of these books.

The text outlines a systematic approach for solving equilibrium problems. It consists of the 11 steps outlined in the text. For a weak acid the first three steps are:

1. List the major species in the solution.
2. Choose the species that can produce H^+ , and write balanced reactions for the reaction producing H^+ .
3. Using the values of the equilibrium constants you have written, decide which equilibrium will dominate in producing H^+ .

In our experience these three steps are the key considerations and should be strongly emphasized. Listing major species gives the student a method for attacking these problems that emphasizes thinking about chemistry rather than memorizing examples. This approach works for all problems involving solution equilibria. Thus the student can look within the problem for the solution rather than outside (memorized pigeonhole examples that usually don't apply to the actual problem at hand anyway). The early introduction to acid-base reactions in Chapter 4 starts the student off with the importance of thinking about reactions. The other important skill is to recognize structural features that give rise to acidic or basic properties. This idea again amplifies the importance of the relationship between structure and properties first covered in Chapter 10. Once the acid (or base) has been identified, a base (or acid) must be found in the species in solution (most often water) that can accept (or donate) a proton. Then it is easy to write a reaction, realizing that acid-base reactions involve a transfer of protons.

The next step is to get values of equilibrium constants for the reaction. This requirement causes difficulty for many students when they must calculate an equilibrium constant from a tabulated one. For example, for the reaction



only K_b for ammonia is given; the student must calculate K_a from

$$K_a K_b = K_w$$

In many textbooks much of the confusion results from calling the above reaction a hydrolysis reaction and giving the equilibrium constant the symbol, K_h . Students think that this is another kind of reaction that must be treated as something different; they get confused and frustrated. We have avoided the use of the term "hydrolysis." There are acids and there are bases. Ammonium ion is an acid because H_3O^+ is being produced. If there are only acids and bases, it is easier for the student to see that it doesn't make sense to use an equilibrium constant that pertains to a base for the dissociation of NH_4^+ . The student is directed to focus attention on a relationship between NH_4^+ and NH_3 , the conjugate acid-base pair. The thought naturally arises to look for a relationship between K_a and K_b for a conjugate acid-base pair. The text emphasizes thinking in a

logical fashion, looking for structural features, and writing the reactions when confronted with acid-base equilibria. This strategy will serve the student much better than attempting to pigeonhole every type of problem.

In the solving of the resulting quadratic equation, the text develops several methods of simplifying the algebra. Assumptions are used to simplify the math when possible, and the method of successive approximations is presented as an approach to use when the usual assumptions are invalid. The method of successive approximation is explained more fully in an appendix and used extensively in the *Solutions Guide*. This tactic further emphasizes the goals of the chapter. We want to deal with the chemistry of acids and bases; we want to make the algebra simple.

The supplements *General Chemistry Problem Solving II* and *Solving Equilibrium Problems with Applications to Qualitative Analysis*, both by Steven S. Zumdahl, can also be used during the coverage of aqueous equilibria. They contain the same systematic approach used to solve equilibrium problems in the main text and provide a great number of simpler, straightforward acid-base and buffer problems. In addition, each chapter contains several multiple-choice questions that the student can use in a self-learning drill type fashion.

Teaching Tips

- Spend time developing the model for acid-base pairs in this section. Students need to be able to look at a chemical equation and identify the acid-base pairs.
- The Bronsted-Lowry definition of acids and bases is preferable because it emphasizes to students that for one reactant to act like an acid the other reactant must act like a base.
- It is important in this section to show students that these substances work together. A proton donor must have a substance to accept the proton. The idea that water can act as both an acid and as a base can be developed here.
- This is the first time students have seen the hydronium ion. Be sure that they understand how it forms from the donation of a proton to water.
- Have students draw molecular level pictures to illustrate the difference between strong acids and weak acids. This idea is important to the understanding of how acids behave.
- We use the terms “ionized” and “dissociated” interchangeably when discussing acids.
- For now, emphasize the competition for a proton between a conjugate base and water in determining which reaction (forward or reverse) dominates.

- After discussing forward and reverse reactions, the students should understand that $\text{C}_2\text{H}_3\text{O}_2^-$ is a better base than H_2O , and H_2O is a better base than Cl^- .
- In discussing the phrase “A weak acid contains a relatively strong conjugate base,” emphasize the term “relatively.” The conjugate base of a weak acid is not a strong base in the way that NaOH is a strong base. This is because for the reaction of the conjugate base in water, the reverse reaction dominates. However, if we compare two unequally weak acids, the conjugate base of the weaker of the two acids is a stronger base, although both bases are considered weak (compared to OH^-).
- Use an equation with two water molecules on the reactant side to show the students that water acts as an acid and as a base. In this way the students can see that a proton is donated from one water molecule and it is accepted by the other water molecule. This allows them to correctly identify acid-base pairs.
- While the students have experience with the concept of an equilibrium constant, they may still not apply it well here. When you discuss the idea that K_w is a constant, be prepared for students who reason “Pure water contains $1.0 \times 10^{-7} \text{ M}$ H^+ and $1.0 \times 10^{-7} \text{ M}$ OH^- . If we add 1.0 mol of HCl to a liter of water, the concentration of H^+ is about 1.0 M. So isn't $K_w = [\text{H}^+][\text{OH}^-] = [1.0 \text{ M}][1.0 \times 10^{-7} \text{ M}] = 1.0 \times 10^{-7}$?” These students do not understand that added H^+ will react with OH^- , thus lowering the concentration of OH^- . If 1.0 mol of H^+ is added to a liter of water, the OH^- concentration would be $1.0 \times 10^{-14} \text{ M}$. This will not be obvious to the students at this point.
- It is helpful for students to go through all equilibrium equations of a polyprotic acid at least once to prove to themselves that only the first dissociation generally matters for the pH.
- Emphasize the definitions of acidic, basic, and neutral solutions. Notice however that the emphasis should be on the relationship of the concentrations of the ions, not on the pH of the solution. This reinforces the concept of *major species* that is highlighted in the text.
- Salt solutions are not new since they consist of conjugate acids or bases in solution. Students may try to treat them differently from weak acids or weak bases, but emphasize the similarities.

CHAPTER FIFTEEN: APPLICATIONS OF AQUEOUS EQUILIBRIA

Chapter Learning Goals:

- | | |
|----------------|--|
| Section One: | To study the effect of a common ion on acid dissociation equilibria. |
| Section Two: | To explain the characteristics of buffered solutions.
To show how to calculate a buffer pH given the concentrations of the buffering materials. |
| Section Three: | To describe the meaning of buffer capacity. |
| Section Four: | To demonstrate how to calculate the pH at any point in an acid-base titration. |
| Section Five: | To explain how acid-base indicators work. |

This chapter is essentially a continuation of Chapter 14. It begins with more complicated acid-base topics: buffers, titrations, and indicators. The problem-solving approach to acid-base problems is continued in this chapter. It helps to emphasize that these problems respond to the same techniques introduced in Chapter 14: writing major species and breaking down problems into smaller parts. An important example of this type of logical thought comes in solving problems that involve complex ion formation. From the values of equilibrium constants, the student knows that the reaction goes almost to completion. Breaking the overall problem down into a stoichiometry problem (complete reaction) followed by an equilibrium problem makes the algebra easier. The chapter emphasizes thinking through the chemistry before doing the algebra. Explicitly calling the first step in the process a stoichiometry problem lets the students see they are not doing anything new. It helps them see why we sometimes approach equilibrium in two steps rather than one. If this explicit label is not provided, it has been our experience that this approach can confuse students. They regard the reaction first going to completion as a mathematical trick rather than chemistry that they have already seen.

Teaching Tips

- A discussion of buffered solutions requires an understanding of LeChatelier's Principle.
- The concept of a buffer is quite complex. It is important to go slowly through the explanation of how a buffer works. Students need to see each of the possibilities and what happens.
- Discuss the chemical properties that are necessary to produce a buffer. Ask questions such as: Why is important to use a weak acid and its salt? Why can't we use a strong acid and its salt?

- The Henderson-Hasselbalch equation assumes that the value of x is negligible in an ICE table. It is not necessarily wrong to assume this (in fact, for most buffers it is fine), but students should always be aware of the assumptions they make.
- Titrations are not “new” but they are good summaries of all of the chemistry that the students have learned in Chapters 14 and 15. Point out Figure 15.3 (the titration of a weak acid with a strong base) to show the students where the solution is buffered, for example.
- Figure 15.4 is a good one to show students that the equivalent point is not dependent on the type of acid or base, but their amounts.

CHAPTER SIXTEEN: SOLUBILITY AND COMPLEX ION EQUILIBRIA

Chapter Learning Goals:

- | | |
|----------------|--|
| Section One: | To show how to calculate the solubility product of a salt given its solubility, and vice versa.
To demonstrate the prediction of relative solubilities from K_{sp} values.
To explain the effect of pH and a common ion on the solubility of a salt. |
| Section Two: | To show how to predict whether precipitation will occur when solutions are mixed.
To describe the use of selective precipitation to separate a mixture of ions in solution. |
| Section Three: | To apply the principles of equilibrium to the formation of complex ions.
To show how complex ion formation can increase the solubility of a salt. |

This chapter provides the first exposure to complex ion formation. Instructors can integrate a discussion of transition metal chemistry (Chapter 21) with the discussion of equilibria.

Some texts treat solubility equilibria before acids and bases. This is a reasonable alternative. It is probably done because for simple salts like AgCl the algebra is easier than that encountered in acid and base problems. This text has not taken such an approach because the acid-base chemistry thoroughly introduces the use of approximations in solving equations, and the solubility of some salts also depends on both acid-base and complex ion equilibria. The foundation is laid in the acid-base sections; all aspects of solubility and complex ion formation can be built upon that foundation. In addition, using only solubility product relationships tends to oversimplify a lot of what goes on in solution. For example, even in the case of silver chloride, a

significant amount of the material in aqueous solution exists as neutral AgCl. People not wishing to tread down this path can rigorously treat acid-base equilibria and then deal with solubility and complex ions in a qualitative fashion.

Teaching Tips

- Students often think of solubility product constants as somehow different from equilibrium constants in Chapter 13 (presumably because of the solids in the chemical equations). Point out the concept of equilibrium is the same in each case.
- Relative solubilities of compounds and their K_{sp} values are only related if there are the same number of ions in the compounds. If there are a different number of ions, we cannot say something like “Because the K_{sp} value of compound A is lower than that of compound B, compound A must be less soluble than compound B.” This claim is true if there is the same number of ions in each compound.
- Have students prove to themselves that if we add chemical equations to get an overall equation we multiply the values of their equilibrium constants to get an overall equilibrium constant.

CHAPTER SEVENTEEN: SPONTANEITY, ENTROPY, AND FREE ENERGY

Chapter Learning Goals:

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|----------------|--|
| Section One: | To define a spontaneous process.
To define entropy in terms of positional probability. |
| Section Two: | To state the second law of thermodynamics in terms of entropy. |
| Section Three: | To discuss the important characteristics of entropy changes in the surroundings.
To apply the relationship between ΔS_{surr} , ΔH , and T (K). |
| Section Four: | To define free energy and relate it to spontaneity. |
| Section Five: | To apply positional probability to chemical reactions.
To relate molecular complexity to entropy. |
| Section Six: | To show how to calculate the standard free energy change in a chemical reaction.
To define standard free energy of formation and show how to use it to predict spontaneity. |
| Section Seven: | To relate free energy to pressure. |
| Section Eight: | To define equilibrium in terms of minimum free energy.
To show how the value of K is related to ΔG° . |
| Section Nine: | To relate work done to the change in free energy. |

The discussion of thermodynamics has been split into two chapters. Chapter 6, Thermochemistry, deals essentially with the first law and conservation of energy. It is placed early so students can get an early exposure to energy relationships, and it gives the background necessary for dealing with the energetics of bonding and phase transitions. Some of this chapter could be discussed at that earlier point. Alternatively, Chapter 6 could be delayed until just before Chapter 17.

Our preference is to keep the chapters split. The students need an early introduction to energy and thermochemistry, but they might be overwhelmed with too much thermodynamics all at once. The later placing of the second law accomplishes this goal. The chapter also places the emphasis of the second law on entropy. This focus would be lost by doing the second law early. The emphasis would tend to shift to free energy.

A focus on entropy, in our opinion, makes the second law clearer. The idea of the formation of stronger bonds (ΔH) serving as a driving force for chemical change can be developed from the topics in Chapter 6. As the intervening chapters are covered, the students become aware that another driving force must be present. Many endothermic processes occur spontaneously. With this appreciation built up over a significant length of course time, it is easier to introduce entropy. Good analogies can be used to qualitatively develop the concept of randomness as a driving force. For example, work has to be done to arrange every student in class in alphabetic order and it work has to be performed to cheat at cards to get a better hand. Chapter 17 is fairly abstract, but it is quite pictorial in approach. We have found that students can develop a better intuitive feel for randomness (entropy) as a driving force for change than they do for free energy. Once the focus of the chapter is placed on entropy, it is easier to accept free energy as a function that accounts for both driving forces for chemical change.

Teaching Tips

- The focus of the section is to introduce the concept that energy naturally changes from useful forms to less useful forms. Each energy transformation results in the same amount of energy but changed to a less useful form. The focus on the quality of energy rather than the quantity of energy is the key idea to present in this chapter.
- The first law of thermodynamics tells us that energy is conserved. The second law gives us some indication of the direction of this energy (the second law is often described as “time’s arrow”). For example, the first law cannot explain why we could not “uncook” an egg by putting a cooked egg in the refrigerator; or why a ball which has just rolled down a hill (and has transferred energy to the hill) could not roll up the hill by spontaneously gathering the newly transferred energy from the hill. The second law informs us that these phenomena are impossible.
- Physical disorder is often given to get the students to think about entropy. Actually, it is best to think of entropy in terms of energy states. However, this is

rather abstract. One way to discuss this is to remind students that molecules have average energies, and temperature is a measure of average kinetic energy. Use Figure 5.21 on page 212 to show students how there is more variance in the “energy spread” of a gas at higher temperatures.

- Students often misinterpret the second law of thermodynamics by thinking that the entropy of the system must increase. It is the entropy of the universe that must increase.
- It is often the case that entropy and enthalpy run counter to each other in the direction of spontaneity. For example, ice requires energy to melt (it is endothermic, and thus the enthalpy term could be said to be unfavorable for spontaneity). However, water is more disordered than ice, so entropy is favorable. Which factor “wins”? It depends on temperature.

CHAPTER EIGHTEEN: ELECTROCHEMISTRY

Chapter Learning Goals:

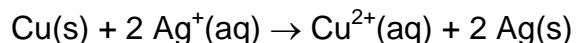
- | | |
|----------------|--|
| Section One: | To describe the half reaction method for balancing oxidation reduction reactions. |
| Section Two: | To review oxidation and reduction.
To define the components of an electrochemical cell.
To distinguish between a galvanic cell and an electrolytic cell.
To define cell potentials. |
| Section Three: | To describe how standard reduction potentials are assigned in terms of the standard hydrogen electrode.
To demonstrate the combination of half-reactions to form the cell reaction.
To characterize a galvanic cell. |
| Section Four: | To relate the maximum cell potential to the free energy difference between cell reactants and products. |
| Section Five: | To discuss the driving force in concentration cells.
To quantify how to calculate the relationship between cell potential and cell concentration.
To show how to calculate equilibrium constants from cell potentials. |
| Section Six: | To discuss the composition and operation of commonly used batteries. |
| Section Seven: | To explain the electrochemical nature of corrosion and describe some means for preventing it. |
| Section Eight: | To describe the stoichiometry of electrolysis reactions.
To show how to predict the order of electrolysis of the |

components of a mixture.
 Section Nine: To discuss the manufacture of aluminum, the chlor-alkali process, and other industrial applications of electrolysis.

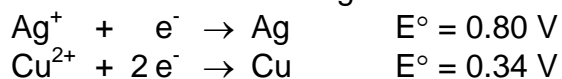
The chapter begins with a review of oxidation-reduction reactions (first discussed in Chapter 4). The emphasis is on using the half reaction method to balance oxidation-reduction reactions.

Free energy is used to develop some of the principles in the electrochemistry chapter, and Chapter 18 follows naturally after Chapter 17. As a pair, these chapters could be placed at different places in the course, as previously mentioned for Chapter 17. Instructors who have not previously covered oxidation-reduction in Chapter 4 need to cover the pertinent topics from Chapter 4 before Chapter 18. At the beginning of the end-of-chapter exercises in Chapter 18, there are problems to review the oxidation-reduction topics covered in Chapter 4.

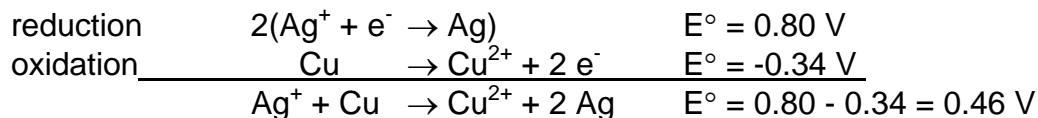
One problem that students have in electrochemistry is dealing with sign conventions for half-cell potentials, and understanding how those half-cell potentials are combined to get an overall cell potential. This process is simplified in this text by emphasizing that, to get an oxidation-reduction reaction from two reduction half-reactions, one of the half-reactions must be reversed (and the sign of the E° value reversed). The decision of which one to reverse is based on getting a net positive potential for the cell. For example, for the reaction



we calculate E° in the following manner. Two half-reactions are



To get a positive E° , the Ag/Ag^+ half-cell involves the reduction (cathode) and the Cu/Cu^{2+} involves oxidation (anode). Thus



This treatment also allows determination of the direction of electron flow without the need for assigning signs to the two electrodes.

A powerful analogy can be drawn in lecture between water and electricity to help the students through the problems of signs. The amount of work done by water falling down a waterfall depends on the amount of water and the height through which the water falls (mgz). The amount of work done by an electric current depends on the amount of charge and the potential difference (nFE). Getting the students to think of

cell potentials as an analogy to the height of a waterfall or as the distance between two points on a continuous scale minimizes sign mistakes. The analogy of charge to a quantity of water tends to minimize the mistake of wanting to use n in calculating cell potentials.

Teaching Tips

- Use the concept of half reactions to emphasize that oxidation cannot occur without reduction, and vice-versa.
- Students need help identifying the half reactions. Writing half reactions is dependent on correctly assigning oxidation states in compounds.
- Be sure that the students balance both atoms and charge in the half reactions. A common error is for students to place the electrons on the wrong side of the half reaction.
- Remind students to equalize the number of electrons transferred. They often forget to do this when they begin to work with these equations. Explain to them that the electrons are transferred and so must be accounted for. This means the number gained must equal the number lost.
- These reactions take time and practice to learn to balance. Students need to pay close attention to the details of the balancing process. Learning to use the method of half reactions on relatively easy equations makes it easier to balance more difficult equations later.
- Explain to students that electrochemical cells are different from ordinary oxidation-reduction reactions because the half reactions are separated from each other.
- Concentration cells can be explained by considering entropy. The greatest entropy exists when the two halves of the cell are mixed to the same concentration. Thus, having differing initial concentrations provides a driving force for the process.
- The lead storage battery is a practical example of electrolysis. Many students know that these batteries are “recharged” when the automobile engine runs. This is a good starting point for a discussion of electrolysis.

CHAPTER NINETEEN: THE NUCLEUS: A CHEMIST'S VIEW

Chapter Learning Goals:

- | | |
|----------------|--|
| Section One: | To relate the stability of a nucleus to the number of protons and neutrons.
To classify the types of radioactive decay. |
| Section Two: | To define and show how to calculate the half-life of a radioactive nuclide. |
| Section Three: | To show how one element may be changed into another using particle bombardment. |
| Section Four: | To discuss radioactivity detection devices.
To show how objects can be dated using radioactive decay. |
| Section Five: | To discuss the thermodynamic stability of the nucleus.
To show how to calculate nuclear bonding energies. |
| Section Six: | To explore the energetics of nuclear fission and nuclear fusion.
To describe how a nuclear reactor works. |
| Section Seven: | To show how radiation damages human tissue. |

Chapter 19 is the only chapter in which the primary focus is on the nucleus. Alternative placement of the topics in this chapter were mentioned earlier: nuclear reactions in Chapter 2, rates of decay in Chapter 12.

There are some important applications of radioactivity. The biological sciences depend heavily on the use of radio-tracers. Analytical techniques, such as radioimmunoassay, use radioactive substances. Some of the more chemical applications of nuclear chemistry are covered in the chapter. The degree to which this chapter is covered will very much depend on the nature of the course being taught.

Teaching Tips

- Balancing nuclear equations provides an excellent opportunity to review balancing by atoms, charge, and the meanings of the mass number and atomic number in element notation.
- The students should be aware that in a nuclear decay the product is a different element. Therefore we cannot balance the equation by balancing the atoms.
- Remind students that in forming ions the number of electrons changes but the number of proton (and therefore the type of atom) remains the same. In nuclear equations if the number of protons is changed the atoms identity is also changed.
- Alchemists attempted to convert base metals (such as lead) to precious metals (such as gold) by chemical means. We now know that this can be accomplished only by nuclear reactions, which is not practical.

- Nuclear transformations are balanced in the same way as nuclear decay reactions.
- Radioisotopes are found in many household devices. Some examples are starters for fluorescent lamps and appliances (promethium-147, krypton-85, and thorium-232); antistatic devices (polonium-210); smoke detectors (americium-241).
- Carbon-14 is produced naturally in the environment. When it is incorporated into carbon dioxide it can be fixed as a carbohydrate in plants. Carbon dioxide containing carbon-14 is an example of a natural source of radioactivity in the environment with which we come in contact.
- The preservation of food by using gamma radiation was first approved by the Food and Drug Administration in 1963. The first foods approved for radiation preservation were wheat flour and potatoes. Not until 1986 could fruits and vegetables be preserved by radiation.
- Table 19.7 gives sources and amounts of radiation exposure for humans. This can be used in a discussion of the natural sources and exposure to radioactivity expected in our lives.
- The use of nuclear imaging techniques is a good starting point for a discussion of the hazards and benefits of nuclear medicine.

CHAPTER TWENTY: THE REPRESENTATIVE ELEMENTS

Chapter Learning Goals:

Section One:	To give an overview of the representative elements. To discuss the relationship of the first member to the rest of the group. To give a general introduction to the commercial preparation of some elements.
Section Two:	To consider the thermodynamics and rates of the reactions of alkali metals with water. To discuss the oxides of alkali metals.
Section Three:	To describe the production and uses of hydrogen. To discuss types of hydrides.
Section Four:	To describe the general chemical properties of the alkaline earth metals. To discuss three-center bonding in beryllium compounds.
Section Five:	To show the general trend from nonmetallic to metallic

	behavior in Group 3A.
Section Six:	To contrast the chemistry of carbon with that of silicon and the other members of Group 4A. To describe carbon oxides. To discuss the characteristics of the +2 and +4 oxidation states of tin and lead.
Section Seven:	To give an overview of the characteristics of the Group 5A elements, with emphasis on bonding.
Section Eight:	To discuss the kinetic and thermodynamic importance of the strength of the triple bond in the nitrogen molecule. To introduce the nitrogen cycle. To describe the structures and uses of major nitrogen-containing compounds.
Section Nine:	To characterize the forms of elemental phosphorus. To describe the preparation and structure of the major compounds of phosphorus.
Section Ten:	To discuss trends in the chemistry of the Group 6A elements.
Section Eleven:	To characterize the forms of elemental oxygen.
Section Twelve:	To characterize the elemental forms of sulfur. To describe the bonding and structure of the oxycompounds of sulfur.
Section Thirteen:	To describe the bonding in compounds of halogens with other nonmetals. To explain the acid strengths of the hydrogen halides. To discuss the chemistry of oxyhalogen compounds. To give the structures of interhalogen compounds.
Section Fourteen:	To describe the uses of the noble gases. To discuss the structures of xenon compounds.

Some examples of how topics from Chapter 20 can be integrated into some earlier chapters have already been discussed. The material in these chapters is organized by family in the periodic table. Many of the end-of-chapter exercises deal with the applications of chemical principles to these elements.

There are several other places in the text where topics from these chapters can be used. There is a short discussion of atmospheric chemistry at the end of Chapter 5, Gases. Instructors choosing to go into greater detail can use material in the chemistry of nitrogen, oxygen, and sulfur. Alternatively, atmospheric chemistry using information can be used to illustrate kinetics. Atmospheric chemistry is controlled to a much greater extent by kinetics than equilibrium. The catalytic destruction of the ozone layer by chlorine from chlorofluorocarbons is an important example that can easily get the students' attention.

Many of the elements that are important in high-technology materials such as

composites, ceramics, photonic devices, semiconductors, and superconductors are representative elements. Instructors can use material from these chapters to illustrate applications of solids in Chapter 10. Because of the discussion of interparticle forces, Chapter 10 is a pivotal chapter in any general chemistry course and is a crossroads for moving into a variety of important applications.

Teaching Tip

- Use this chapter as a good summary of many of the concepts discussed in other chapters. For example, Section 20.1 requires the students to recall Section 7.12.

CHAPTER TWENTY-ONE: TRANSITION METALS AND COORDINATION CHEMISTRY**Chapter Learning Goals:**

- | | |
|----------------|---|
| Section One: | To discuss electron configurations and general trends in the properties of the first-row transition metals.
To introduce the 4d and 5d transition series. |
| Section Two: | To discuss some chemistry of the 3d transition metals. |
| Section Three: | To define terms used in coordination chemistry.
To discuss some common chelates.
To give rules for naming coordination compounds. |
| Section Four: | To introduce and illustrate types of structural isomerism and stereoisomerism.
To relate molecular structure to optical activity. |
| Section Five: | To use the localized electron model to explain the interactions between a metal ion and ligands. |
| Section Six: | To describe the crystal field model and use it to explain the magnetism and colors of coordination complexes. |
| Section Seven: | To describe the function and structures of some biologically important coordination complexes. |
| Section Eight: | To describe processes for concentrating minerals in ores.
To discuss the major processes in hydrometallurgy.
To describe the blast furnace, open hearth, and basic oxygen processes for steelmaking.
To discuss the thermodynamics of steelmaking. |

We have previously mentioned that material from this chapter can be used in conjunction with a discussion of metallurgy (Chapter 10) or complex ions (Chapter 16). Another alternative is to cover transition metal chemistry after Chapter 9 to illustrate bonding models.

Two further points should be made concerning this chapter. First, the only discussion of

electron configurations of transition metal ions appears in Section 21.1. Only the electron configurations of ions of representative elements are considered in the discussion of ionic bonding in Chapter 8. Second, the concept of isomerism is introduced in this chapter. Structural isomerism and stereoisomerism are illustrated with examples of transition metal complexes. Instructors who wish to introduce isomerism during earlier coverage of bonding may want to use Section 21.4. Alternatively, organic examples can be used for an early introduction to isomers. They are covered in Chapter 22 (structural and geometrical isomerism). Optical isomerism and chirality in organic compounds are introduced in Section 22.6 (carbohydrates).

Teaching Tips

- A compound such as $[\text{Co}(\text{NH}_3)_5\text{Cl}]\text{Cl}_2$ requires the students to have an understanding of polar covalent bonds (in the NH_3), coordinate covalent bonds (between the NH_3 and Cl^- ligands and the metal ion), and ionic bonding (between the chloride ions and the complex ion).
- It is helpful for many students to make physical models of the different types of isomers. This is especially true for the optical isomers with bidentate ligands such as shown in Figure 21.16.
- The crystal field model provides another opportunity to discuss the nature of science and the use and purpose of models in science. For example, we begin by defining ligands as negative point charges (even though this is not technically true). We use this definition when developing the model.
- Figure 21.21 shows a problem with the ligand as negative point charge definition. That is, all d-orbitals increase in energy. From discussion in Chapter 6 and Chapter 17, the students should realize that this should not occur. This shows us that the crystal field model, while correctly discussing splitting (and thus the color of such compounds), is a poor model for discussing energy.
- Consider using the color wheel to discuss complementary colors. Thus, red is complementary to green, orange is complementary to blue, and yellow is complementary to violet.

CHAPTER TWENTY-TWO: ORGANIC AND BIOLOGICAL MOLECULES

Chapter Learning Goals:

- | | |
|----------------|---|
| Section One: | To describe the nomenclature system for organic chemistry.
To discuss isomerism in organic molecules.
To give the rules for naming alkanes. |
| Section Two: | To introduce unsaturated hydrocarbons.
To discuss isomerism in alkenes.
To describe common reactions involving multiple carbon-carbon bonds. |
| Section Three: | To discuss the structures and reactivities of aromatic hydrocarbons. |
| Section Four: | To introduce the basic functional groups and give a few characteristics of each one. |
| Section Five: | To show how polymers are formed from monomers via two types of polymerization reactions. |
| Section Six: | To describe the levels of structure and the functions of proteins.
To describe the structures and isomerism of simple carbohydrates.
To describe the structures and functions of polymers of glucose.
To describe the structures and functions of nucleic acids. |

Chapter 22 emphasizes the structures of several classes of organic compounds. Instructors can use this chapter to lead to several applications. Energy and fuels can be emphasized by following up a discussion of hydrocarbons with petrochemicals (Section 22.4).

The structures and properties of synthetic polymers can be covered by emphasizing Section 22.5. Organic chemistry can be introduced after a discussion of bonding or, particularly if synthetic polymers are stressed, as an illustration of using interparticular forces in correlating structure and properties.

Optical isomerism in organic compounds and the chiral carbon atom are introduced in the section on carbohydrates (22.6). There are end-of-chapter problems on optical activity that deal only with simpler organic compounds for use by instructors who wish to introduce the topic to students without going into a detailed discussion of carbohydrates.

Some of the end-of-chapter exercises on nucleic acids deal with the genetic code and errors in the transmission of genetic information. The information science aspects of nucleic acids are an interesting application of entropy.

Teaching Tips

- To introduce the topic of organic chemistry and its impact on our lives ask the students what the classroom would look like without carbon containing compounds. You will need to assist them in figuring out what things contain carbon, but it will not be long before they can appreciate how many of the objects surrounding us are organic compounds.
- You will find a good set of three-dimensional models very helpful in presenting this material. Use the models to illustrate the shapes of the molecules discussed in this section. The tetrahedral shape of methane is difficult to visualize without a model. Inexpensive classroom sized models can be built from painted Styrofoam balls and wooden dowels.
- Strings of children's pop beads can be used to illustrate that bending or twisting the string does not produce a change in the structure. To create a different order of beads you must take apart the chain and put it together differently.
- Organic nomenclature can seem overwhelming at first. It is not possible to memorize all the names and formulas for organic molecules. Students need to learn the names for the basic root and substituents. Once these are learned naming organic compounds becomes a problem-solving activity.
- Be sure to include some examples where the longest carbon chain is not horizontal. Students find it difficult to understand that a carbon chain with several bends in it is not branched. It is helpful to tell the students that the longest chain is the one which they can follow without lifting their pencil along the carbons. Circling or highlighting the longest chain of carbons can help when you are teaching the students to identify the substituents.
- It is helpful to show students how to think about isomer problems. An organized approach for drawing and naming isomers makes the process much simpler. Start with the longest carbon chain and remove one methyl group at a time and try to place it on as many new carbons as possible without producing duplicates. Naming each molecule will help to eliminate duplicates. If the names are the same then the molecules are the same even if they look different in the drawing.
- In substitutions reactions be sure to remind the students that any one of the hydrogens can be substituted. Not all substitutions will produce different products. This can be confusing to some students. Use three-dimensional models to illustrate this one molecule at a time.
- A model that shows the difference between cyclohexane (a puckered ring) and benzene (a flat ring) would be useful in helping students understand aromatic hydrocarbons.

- You may wish to discuss the delocalization of the electrons in the benzene ring as an example of how our bonding model is limited. Drawing the benzene ring structure with the circle inside the hexagon reminds us that benzene behaves differently than a carbon molecule with a double bond.
- When determining whether an alcohol is primary, secondary, or tertiary the focus should be on the carbon where the -OH is attached. Students often focus on the -OH group itself instead and this does not tell anything about the classification.
- Section 22.6 is an excellent place to review what causes a side group to be polar or nonpolar. Refer the students back to the ideas they learned in bonding and work with them to determine why each of the side chains are classified as polar or nonpolar.
- The term “hydrophobic” is actually a misnomer. While the non-polar side chains do not mix well with water, they do not repel the water. Water molecules are more attracted to other water molecules than to the non-polar side chains.
- An excellent model of secondary structure is the coiled cord from a telephone. Another possibility is to use the children’s pop beads and twist them into a helix to represent secondary structure while showing the primary structure by the order of the pop beads. A folded piece of paper can represent the pleated sheet secondary structure.
- You can use the analogy of a ladder to help students understand the structure of DNA. The sides of the ladder represent the alternating sugar and phosphate groups and the rungs of the ladder represent the base pairs that are hydrogen bonded together.
- Fats containing more unsaturated fatty acids are liquid at room temperature.
- Lipids are different than most other organic and biomolecules because they are distinguished not by a particular functional group but by their solubility in nonpolar solvents.

PART III: SUPPLEMENTARY MATERIALS

FOR THE INSTRUCTOR

Supporting instructor materials are available to qualified adopters. Please consult your local Cengage Learning, Brooks/Cole representative for details. Visit www.cengage.com/chemistry/zumdahl to:

- See samples of materials
- Request a desk copy
- Locate your local representative
- Download electronic files of the *Instructor's Resource Guide*, *Instructor's Resource Guide for Experimental Chemistry*, and other helpful materials for instructors and students.

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- Image libraries in PowerPoint and JPEG formats that contain electronic files for all text art, most photographs, and all numbered tables in the text. These files can be used to create your own transparencies or PowerPoint lectures.
- JoinIn "clicker" Slides: Join "clicker" questions written specifically for the use of Chemistry with the classroom response system of your choice that allows you to seamlessly display student answers.
- PowerPoint slides to accompany the Inquiry Based Learning Guide (see Inquiry Based Learning Guide description below).
- **Complete Solutions Manual** (Thomas J. Hummel, University of Illinois, Urbana-Champaign) The *Complete Solutions Manual* contains detailed solutions to *all* of the end-of-chapter problems with the exception of the *For Review Questions*, available on the student website, and a discussion of the *Active Learning Questions* as found in the online *Instructor's Resource Guide*. This supplement is intended for the teacher's convenience.
- Sample chapters from the *Student Solutions Manual* and *Study Guide*.
- *Instructor's Resource Guide* for the textbook and for *Experimental Chemistry*.
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- Ensure that each student gets a different version of the problem by selecting from preprogrammed algorithmic questions.
- Edit or author algorithmic or static questions that integrate into the existing bank, becoming part of the question database for future use.
- Choose problems designated as single-skill (easy), multi-skill (medium), and challenging and multi-skill (hard).
- Customize tests to assess the specific content from the text.
- Create several forms of the same test where questions and answers are scrambled.



OWL: Online Web-based Learning by Roberta Day and Beatrice Botch of the University of Massachusetts, Amherst, and William Vining of the State University of New York at Oneonta.

Developed at the University of Massachusetts, Amherst, and class tested by tens of thousands of chemistry students, OWL is a fully customizable and flexible web-based learning system. OWL supports mastery learning and offers numerical, chemical, and contextual parameterization to produce thousands of problems correlated to this text. The OWL system also features a database of simulations, tutorials, and exercises, as well as end-of-chapter problems from the text. In addition, OWL now includes ChemWork assignments, which help students learn key chemical concepts while guiding them in the process to become problem solvers. (See description below.) With OWL, you get the most widely-used online learning system available for chemistry with unsurpassed reliability and dedicated training and support. Also new in OWL is Go Chemistry™—27 mini video lectures covering key chemistry concepts that student can view onscreen or download to their portable video player to study on the go! The optional **e-Book in OWL** includes the complete electronic version of the text, fully integrated and linked to OWL homework problems. Most e-Books in OWL are interactive and offer highlighting, notetaking, and bookmarking features that can all be saved. To view an OWL demo and for more information, visit www.cengage.com/owl or contact your Cengage Learning Brooks/Cole representative.

ChemWork

ChemWork assignments offer students another opportunity to practice. These problems are designed to be used in one of two ways: the student can use ChemWork to *learn* the problem solving process (while doing actual homework problems) or the students can use ChemWork as a *capstone* assignment to determine whether they understand how to solve problems (perhaps in final preparation for an exam).

ChemWork assignments test students' understanding of core concepts from each chapter. If a student can solve a particular problem with no assistance, he/she can proceed directly to the answer and receive congratulations. However, if a student needs help, assistance is available through a series of hints. The procedure for assisting the student is modeled after the way an instructor would help a student with a homework problem in his or her office. The hints are usually in the form of interactive questions that guide the student through the problem solving process. Students cannot receive the right answer from the system, rather it encourages them to continue working on the problem through this system of multiple hints. *ChemWork* is chemically and numerically parameterized so that each student in the course receives a unique set of problems.

Enhanced WebAssign

Enhanced WebAssign, a robust, easy-to-use online learning system, includes algorithmic textbook problems with rich media learning resources, such as *ChemWork* assignments, tutorials, animations, and videos. Instructors can create assignments from a ready-to-use database of textbook questions or instructors can write and customize their own exercises with results recorded in its corresponding gradebook. To learn more, go to www.webassign.net/brookscole.

Instructor's Resource Guide (Donald J DeCoste, University of Illinois, Urbana-Champaign) Available on the instructor companion site and on PowerLecture, this downloadable manual includes suggestions for alternative ordering of topics, lesson plans of media resources correlated to section, a section of notes for teaching assistants, suggested solutions for the *Active Learning Questions* from the textbook, and amplification of strategies used in various chapters.

Instructor's Resource Guide for Experimental Chemistry (James Hall, University of Massachusetts, Lowell) Available on the instructor companion site and on PowerLecture, this PDF manual contains tips on running experiments, approximate times for each experiment, and answers to all prelab and postlab questions posed in the laboratory guide.

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Visit the *Chemistry* website at www.cengage.com/chemistry/zumdahl to see samples of select student supplements. Students can purchase any Cengage Learning product at your local college store or at our preferred online store www.ichapters.com.

Student Companion Website

Accessible from www.cengage.com/chemistry/zumdahl, this site provides online study tools including practice tests, flashcards, Chemical Connections, the answers to the *For Review* questions from the textbook, and Careers in Chemistry.



Go Chemistry™ for General Chemistry 27-Module Set ISBN-10: 0-495-38228-0; ISBN-13: 978-0-495-38228-7

Go Chemistry™ is a set of easy-to-use essential videos that can be downloaded to your video iPod or portable video player—ideal for the student on the go! Developed by award-winning chemists, these new electronic tools are designed to help students quickly review essential chemistry topics. Mini video lectures include animations and problems for a quick summary of key concepts. Selected Go Chemistry modules have e-flashcards to briefly introduce a key concept and then test student understanding of the basics with a series of questions. Go Chemistry also plays on QuickTime, iTunes, and iPhones. Modules are also available separately. To purchase, enter ISBN 0-495-38228-0 at www.ichapters.com.



OWL for General Chemistry

See the above description in the instructor support materials section.

Study Guide (Paul B. Kelter, Northern Illinois University); ISBN-10: 0-547-16872-1; ISBN-978-0-547-16872-2

Written to be a self-study aid for students, this guide includes alternate strategies for solving problems, supplemental explanations for the most difficult material, and self-tests. There are approximately 500 worked examples and 1200 practice problems (with answers), designed to give students mastery and confidence.

Student Solutions Manual (Thomas J. Hummel, University of Illinois, Urbana-Champaign) ISBN-10: 0-547-16856-X; ISBN-13: 978-0-547-16856-2

This manual provides detailed solutions for half of the end-of-chapter exercises (designated by the blue question numbers) using the strategies emphasized in the text. This supplement has been thoroughly checked for precision and accuracy. Answers to the *For Review* questions can be found on the student website.

Experimental Chemistry (James Hall, University of Massachusetts, Lowell)

ISBN-10: 0-547-16854-3; ISBN-13: 978-0-547-16854-8

This lab manual provides an extensively revised laboratory program compatible with the text. The 48 experiments present a wide variety of chemistry, and many experiments offer choices of procedures. Safety is strongly emphasized throughout the program.

Inquiry Based Learning Guide (Donald J. DeCoste, University of Illinois, Urbana-Champaign) ISBN-10: 0-547-16871-3; ISBN-13: 978-0-547-16871-5

The questions and activities in the *Inquiry Based Learning Guide (IBLG)* are designed to get students to consider the underlying concepts involved in understanding chemistry—to "think like chemists" and develop a richer, more dynamic understanding of the content. By being engaged in the material and needing to explain how they apply it rather than just calculating an answer, students are given an opportunity to strengthen their problem solving abilities, and in turn teachers are given an opportunity to identify students' misconceptions that may not be apparent in their answers to traditional, mathematically-based questions.

Essential Algebra for Chemistry Students, Second Edition by David W. Ball, Cleveland State University. ISBN-10: 0-495-01327-7; ISBN-13 978-0-495-01327-3

This short book is intended for students who lack confidence and/or competency in their essential mathematics skills necessary to survive in general chemistry. Each chapter focuses on a specific type of skill and has worked-out examples to show how these skills translate to chemical problem solving. Includes references to OWL, our Web-based tutorial program, offering students access to online algebra skills exercises.

Survival Guide for General Chemistry with Math Review and Proficiency Questions, Second Edition by Charles H. Atwood, University of Georgia. ISBN-10: 0-495-38751-7, ISBN-13 978-0-495-38751-0

Intended to help students practice for exams, this survival guide shows how to solve difficult problems by dissecting them into manageable chunks. The guide includes three levels of proficiency questions—A, B, and minimal—to quickly build confidence as students master the knowledge needed to be successful in their course.

FOR THE LABORATORY

Cengage Learning Brooks/Cole Lab Manuals. We offer a variety of printed manuals to meet all your general chemistry laboratory needs. Instructors can visit the chemistry site at www.cengage.com/chemistry for a full listing and description of these laboratory manuals and laboratory notebooks. All Cengage Learning lab manuals can be customized for your specific needs.

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PART IV: BEING AN EFFECTIVE TEACHING ASSISTANT

by Steven Zumdahl

In large general chemistry programs, teaching assistants typically play a crucial role. Their attitudes and levels of preparation greatly affect the quality of education that the students receive. Following is an excerpt from the TA manual used at the University of Illinois. We hope these comments will also be helpful to TAs at other institutions.

INTRODUCTION

Teaching assistants form the heart of any large general chemistry program. The quality of the instruction is highly dependent on your attitude and your teaching skills. You have most of the close contacts with the students and thus can most directly influence their performance. It is very important that you take this responsibility seriously and do the best job possible. Following are some specific suggestions that you should find helpful.

GENERAL COMMENTS

1. **Spend some time thinking about the learning process.** Although you may never have taken a course in the psychology of education, you have a great deal of experience in the field—you have been a student for over 15 years. Think about your experience as a student. Figure out why some teachers were very effective and others very ineffective. Your long experience as a student is excellent preparation for your teaching career.

2. **Your attitude is crucial** in setting the correct tone for your class.

Be enthusiastic. Project a positive attitude toward your students and the subject. Presumably you are excited about chemistry or you wouldn't be pursuing a graduate degree in the subject. Let this enthusiasm shine through.

Avoid being cynical and negative. Poor attitudes will rub off on your students. If you have a criticism of the way the course is being run, see the professor for the course—do not gripe to the students. Your attitude has a very important effect on the attitudes of your students. Setting a good example is one of the most important things a good teacher does for his/her students.

3. **Have appropriate expectations** of your students.

Do not expect too little from your students. If you think of your students as dullards who have no hope of really learning much chemistry, this will come through to them. Be positive, treat them like the talented people they are, and encourage them to live up to their potentials. This may sound corny, but is very important. On the other hand, recognize that the students in your class are not graduate students. Don't set unreasonable goals for your students.

4. **Be sensitive and even-handed.** Treat your students with respect. Remember they are quite heterogeneous in background and experience. Humor in the classroom is very refreshing, but it must be used appropriately. Don't be offensive, and if you direct humor at anyone, it is best to direct it toward yourself.

Do not flaunt your authority. You can be firm and maintain high standards without being dictatorial. Do not do anything that might be even remotely interpreted as sexual harassment. This will not be tolerated.

Treat everyone the same. If you have problems with a particular student, discuss the problem privately.

Always remember that you are now part of the teaching staff, not a student (at least not in this program). Be sensitive and sympathetic toward the students, but do not try to be like another friend of the students. You need to keep some separation between yourself and the students to keep objectivity.

5. **Talk to the students about educational goals.** Chemistry provides excellent fringe benefits. To solve a problem in chemistry, a student must be able to think critically and creatively and must be careful and logical. Problem solving is a skill that is highly useful to anyone—scientist, attorney, doctor, laborer, and so on. Many curricula that require chemistry do so not only because chemical knowledge is important, but because chemistry as a discipline provides an excellent vehicle for learning to “think.” You should be aware of this as you teach and should point it out to your students often. Teach in a way that encourages students to be critical and creative.

Above all, set a good example about what it means to be educated, and let the students know you are genuinely interested in helping them improve.

6. **Help your students learn to be flexible.** Chemistry is difficult for beginning students mainly because it is complicated—chemical systems are typically influenced by a large number of factors. Because of this, a fail-safe method is to learn not to be locked into a particular method, but to let the problem guide them toward its solution. To help them learn to do this, you should stress flexibility in solving problems. Use a variety of methods in your teaching. Above all, stress the chemical principles.

Most students in general chemistry will never have to solve a stoichiometry or gas law problem in connection with their careers, but an **understanding** of chemical principles may well be important to them, and the ability to solve a variety of problems will certainly be important to them. Thus you must be careful to test the students' understanding; do not make simply getting the right answer to a given problem the only goal. Give some "algorithm busters" on your quizzes to prevent students from simply memorizing methods for solving particular problems. This makes teaching more difficult and may result in some tension between you and the class as you challenge your students to stretch themselves, but it is the only honest way to teach.

7. **Do not quarrel with the course materials.** You may not like the text, lab book, or other resources, but don't tell the students. If you destroy their faith in the only learning materials available to them, you will seriously undermine their morale. You are welcome, even encouraged, to be critical, but pass along your criticisms to the course professor, not the students.

SPECIFIC TEACHING TIPS

1. **First impressions are important.** On the day you meet your class for the first time, you may be nervous and uncertain, but project a positive attitude to your students. After the first day of classes in the fall semester, we sometimes have panicky students come to us with comments such as "The TA doesn't even know how the course grade is determined," or "The TA said 'I wanted a lab assignment but they gave me a recitation section even though I can't talk to groups,'" or "The TA said, 'I used this text in college and it's terrible.'" A little honest humility is okay, but don't overdo it. The students expect competence and confidence on the part of their teacher. It is very important to be especially well prepared on your first day of class.

The first day you meet your class

- a) Write the following on the board:

the course and section
your name
your office and office hours
- b) Introduce yourself. You might tell the students a bit about yourself, including your interests and your intended career.
- c) An important function of the first day of class is to clearly indicate how the discussion group fits in with the rest of the course and what you expect of them in your class section.

- d) Give the students specific information about factors that affect their grades: attendance, test and quiz policies, homework, your expectations, the grading system, and so on.

2. **Be punctual. Be prepared.**

Not even the most experienced and skillful teachers try to teach without preparation. **DO** the assigned problems. A problem may look simple and straightforward, but there are few moments more embarrassing than being unable to do an assigned problem when you are standing in front of a class. Some problems have quirks in them that are easily missed in casual appraisal. Do the problems before going to class. Do not assume you can teach successfully by “shooting from the hip.”

- 3. **Dress casually but neatly.** Dress as if you care about the class and have prepared for it. You must **appear** to be in charge of the class if you hope to maintain effective control in the classroom.
- 4. **Wear a watch.** Most of the classrooms do not have clocks. If you are going to use time effectively, you must know how much time remains in the class period.
- 5. **Learn your students' names** as quickly as possible. Call on the students by name.
- 6. **Speak clearly and loudly and write legibly on the board.** Make sure your voice can be heard and your writing read in the back of the room.
- 7. **The quiz/recitation class should be a discussion, not a lecture.** The primary purpose of the recitation section is to provide an opportunity for discussion. Encourage all students to participate. Be careful not to let one or two students dominate every discussion. To teach effectively, it is essential to get feedback from your class concerning what they need from you. Encourage discussion by asking effective questions. A simple “any questions?” is usually not sufficient. Refer to specific material and call on students by name.

When a question arises for which you do not know the answer, say, “I don’t know,” and find the answer before the next quiz session. Do not try to bluff. You will look bad and, even worse, the students will get incorrect information.

To be an effective discussion leader, you must prepare before class. Questions that are at the right level and that are interesting and challenging will draw students into the discussion. Quite often, once students are involved in a discussion, you need only to interject a few “mid-course” corrections to guide the students through the material. These may include content presentations, problem-solving activities, or additional questions.

Come to class with a few relatively challenging conceptual questions. Ideally these questions will require the students to interact with new material and cannot be solved by applying a simple algorithm or by just using memorized answers (like rules for naming compounds).

Thus one important goal is to get the students active in class. Another is to get the students to think about chemical systems and WHY chemists solve problems the way they do instead of just HOW to solve them. We can show the students how to solve problems over and over but because this approach does not force them to directly interact with the concepts, it often does not “take.” Students are sometimes more critical of each other’s ideas than we are and they will be more likely to ask other students, “How did you get that?” or “How did you know how to do that?” When this happens, more learning occurs. To create such an environment requires that you set up a friendly atmosphere in which the students know that while their ideas will be critiqued, they should not take this criticism personally. Use at least some of the first day of class to explicitly discuss these ideas with students and let them know that they will be responsible for much of the interaction in the section.

8. **Don’t be too subtle.** Many freshmen are used to being spoon-fed, and they need considerable direction. In your presentation, be very clear about what you are doing and why. This is especially important when doing problems. There is a natural tendency to rush to the points about the problem that you want to emphasize and, as a result, to inadequately describe the steps in the problem. As you discuss a solution to a problem, write pertinent information on the board and label everything carefully.

It is important to recognize the differences between the spoken and the written word. Writing should be elegant and concise, but it will be read again and again. Verbal teaching requires organization, simplicity, and repetition. Don’t be afraid to be simple and direct. Make sure students understand which concepts are crucial. Freshmen tend to be pragmatic about the course. They have heard that chemistry is hard and they want to know what will help them succeed (i.e., get good grades on the exams). You have to help them with this. They don’t know what is critical and what is not. Point out, during the course of the discussion, things that you expect will appear on exams.

9. **Effective quizzes.** Your quizzes should be short (~15 minutes) and should emphasize the material that will be tested on the exams. You should have a range of easy to hard questions. If your quizzes are too easy, your students will suffer on exams. If they are too hard, the students will become frustrated. Grade quizzes promptly. When you return the quiz, ask for questions or go over the entire quiz if you think it is necessary. At minimum, make sure everyone knows the correct answers to all of the questions.

ASKING GOOD QUESTIONS—HOW TO ENCOURAGE DISCUSSION

There are many situations where a teaching assistant must be able to ask and answer questions skillfully. Since much of the time spent in recitation/quiz sections will be in a question-answer format, we want to help you become a more effective discussion leader. The idea is to cover the material and involve the students in the process. If a TA simply lectures in the recitation section, the material will be covered, but student interest may be low. On the other hand, a TA standing in front of the class fielding student questions might keep the students more involved, but may not cover the material adequately.

An alternative is for the TA to ask questions of the students. Questions that are at the right level and that are interesting and challenging will draw the students into the discussion. Quite often, once students are involved in a discussion, you need only interject a few “mid-course” corrections to help guide the students through the material. The mid-course corrections may consist of mini-lectures, problem solving, or additional questions.

The job of discussion leader is not easy. It requires thought and preparation. You must consider how to start a discussion and how to maintain the discussion.

Starting a discussion—“initiating questions.” An initiating question is a good way to begin a discussion of a particular topic. To be effective, the questions must be geared to the course content and your best estimation of the ability of the class. These questions should be planned in advance and can be arranged in order of increasing difficulty or abstraction. A question that requires only rote memory might be first, followed by one that tests for understanding or ability to apply rote knowledge.

Choosing the type of question you ask can be crucial because it can set the tone of the discussion. For example, consider the following question in two different forms (A and B):

- A. What is lattice energy and why is it important?
- B. Given that the lattice energy of the ionic compound X is ____ and that of the ionic compound Y is ____, what does this tell you about the relative stability of the two compounds and why? The relative ease of their formation and why? Why are the values negative?

The A version of the question could allow for a textbook definition from a student, which would require more questioning by you to try to determine if the students understand the concept of lattice energy. If the textbook definition is accepted as the answer, the students may come away thinking that memorizing is equivalent to understanding. The question in form B requires that the students think about what lattice energy means (thus they have to essentially answer form A anyway!) and how to apply this knowledge. Asking the question in form B instead of A also signifies that you are not interested in definitions from the students but in higher-level thinking. You could always ask the question in form A later if the students are having difficulty.

Brief descriptions of different levels of questions, as well as sample questions, are given below:

1. **Knowledge**—Simple recall of facts, rote memory.
Examples:
 - What is the first law of thermodynamics?
 - What is the symbol for work? heat?
2. **Comprehension**—Ability to translate into one's own words: using a given equation to solve a problem; translating a literal statement into an equation.
Example:
 - Give an equation showing that pressure depends inversely on the volume of 1 mole of an ideal gas at 25°C.
3. **Application**—Ability to apply concepts to a specific situation; recognizing and solving a problem where the equations are not given.
 - Give an example of a process that demonstrates the first law of thermodynamics.
 - If an ideal gas expands reversibly and isothermally, will ΔS be positive or negative?
4. **Synthesis**—Students must assemble components into a form that is new to them.
Example:
 - What factors would you have to take into consideration if you were designing a high-efficiency engine?

Low-level questions such as knowledge questions are usually easy to construct. Moreover, almost any student who is reasonably well prepared should be able to answer them. These types of questions usually are short, have predictable responses, and will reinforce student participation. Higher-level questions, such as the synthesis type, usually generate greater student interest, but they have many possible answers. You must be ready for a whole range of responses and be able to move ahead from that point. This is not an easy task and will take practice.

Students will often answer questions in ways the TA is not expecting. A student may have misunderstood the question, answered incorrectly, or wanted to interject something relevant to the topic being discussed. To help you minimize awkward pauses while you try to digest a student response, a few pointers are outlined below on how to follow up a student response.

Follow-up questions. Often a student's response is not of high quality (i.e., there may be some ambiguity, or the response may be incomplete). Sometimes there will be no response at all! If a student's response is unclear, or if it is important that a point be emphasized, have the student clarify and/or elaborate on his or her previous statement. For example:

- What do you mean by "exothermic"?
- Can you give me an example of that idea?
- Can you be a little more specific?

A student who gives a high-quality response can be used to teach the rest of the class. In this way, the TA does not monopolize the discussion. Also, the student who gives the response will be reinforced, which will increase the likelihood that he or she will participate in future discussions. It is especially important to follow up when a student gives a wrong answer. Usually when one student is wrong, several others suffer the same misunderstanding. Thus when one person is corrected, several may benefit. Be sure to be tactful in correcting students. Sometimes it is even useful to challenge a correct answer and make the student defend his or her stand.

An effective method to bring other students into the discussion is to ask another student to comment on a particular response:

- Mary, can you clarify what Dave means by the term spontaneous?
- Jim, is that correct?

This type of follow-up question is especially useful if the student gives an incorrect response. Instead of your correcting the student, a peer can get involved. Also, if two students have different points of view, an impromptu discussion may ensue that incorporates material relevant to the class. These discussions often stimulate interest; students are likely to pay attention. You should intercede when necessary to refocus the discussion and to emphasize certain points, or to help clarify various students' positions. Always summarize the discussion for the students.

When a student does not respond at all, do not automatically move on to someone else. There are ways to coax an answer from the student. The question can be rephrased, or it can be broken down into simpler parts. You can also give the student hints, such as pointing out certain parts of the chalkboard that have relevant information. Remember, your goal is to get students into the discussion, not to embarrass them. Good judgment is needed in deciding whether or how far to probe a student's silence.

Summarizing the discussion. Although there are many advantages to the type of discussion outlined above, one disadvantage is that the information on a particular topic may accumulate in a rather disjointed fashion. It is very important at the end of a question-answer discussion that you carefully summarize important points. This serves as a useful review for the students, focuses their attention on the most important points of the discussion, and helps them organize the information to which they have been exposed.

HINTS FOR SUCCESSFULLY LEADING A DISCUSSION

Ask for questions with confidence and sincerity. Your manner should imply that you really want questions. You always take a chance by opening up for discussion (you may not know all the answers) but it is worth the risk. Before class, think about the questions that may come up and go through your answers. If a question is asked that you can't answer, you might initially seek help from the class. However, if things are not clarified rapidly, say that you will bring the answer to the next class and do it. In the meantime, let your students know that you expect answers to your questions by asking meaningful questions confidently and openly.

Avoid trite questions. Questions about things that are obvious are best left out of a discussion. Don't insult the intelligence of your students! Other types of questions that do not promote discussion are, "Right?", "Any questions?", and questions that can be answered by a simple "Yes" or "No."

Wait! Wait a few seconds before calling on someone for a response. This will give everyone a chance to think about the answer. If students are asked to respond too quickly, they may freeze and not be able to think of the answer. Waiting also indicates to the students that you really expect a thoughtful response.

Select specific respondents. Call on students by name. This will allow more people to contribute to the discussion and will prevent eager-beavers from monopolizing the discussion.

Distribute questions. Feel free to call on anyone in the class, including those who do not have their hands up. Shy students may never involve themselves in a discussion unless they are prompted. Furthermore, if students know they might be called on, they tend to spend more time in preparation for class.

Listen to the response. Often if you have a certain answer in mind, you may not recognize a different but equally correct response. Remember that there are many ways to approach problems.

Reinforce responses. Let students know that you are pleased by their participation. Words like "good" or "excellent" are appreciated by students. You can also provide reinforcement by smiling and/or nodding agreement as the student responds. Even for wrong answers, try to reinforce the act of responding.

Follow-up questions for wrong responses at least let the student know that you are concerned.

Use the students. Students can sometimes explain complex phenomena or problems in novel ways. They may do a better job than a TA of explaining material on the level of their peers.

Discourage guessing. Make it clear to your students that the thought process they use is more important than the answer. If a student is guessing, a follow-up question should make it obvious.

In summary. The question-answer format has proven to be an effective way to present material in recitation sections. Among the many advantages of this format are that it

- allows you to assess where the class needs help,
- increases student interest,
- stimulates student thinking,
- allows student contributions,
- provides emphasis and reinforcement of important points,
- adjusts instruction to the level of the class,
- checks the effectiveness of the instruction, and
- reveals student attitudes.

It is also fun to teach this way.

QUESTIONNAIRE FOR STUDENT FEEDBACK (Copy follows)

A questionnaire has been designed that will provide you with some feedback about your teaching. Only you will see the results of this questionnaire. The best time to use the questionnaire is the last 15-20 minutes of class about 4 weeks into the course.

Introduce the questionnaire in a serious manner, emphasizing positive aspects (e.g., how the students can help you do a better job of teaching).

In interpreting the results, remember that students can be quite critical and you might not receive much praise. Look for patterns in the results. For example, if most students feel your explanations are vague, they probably are. You should think more carefully before class about the concepts you will be discussing. Be receptive to constructive criticism, but keep criticism in perspective. Students are generally quite fair but sometimes they look for someone else to blame for poor performance. If your students' evaluations of your performance are very harsh, see the professor for your course for help as soon as possible. Anger at the class or depression on your part will only aggravate any problems you may have. It is far better to regard student evaluations as raw data (of varying quality) that may help you develop better teaching skills.

GENERAL CHEMISTRY MIDSEMESTER STUDENT QUESTIONNAIRE

To the Student: This questionnaire is designed to provide feedback to your TA to help him/her do a more effective job. Please fill it out thoughtfully and honestly. You may write comments in response to any question and at the end if you wish.

- | | | | | | | | | |
|-----|---|--------------------|---|---|---|---|---|----------------------|
| 1. | Is the instructor prepared for class? | Always | 1 | 2 | 3 | 4 | 5 | Never |
| 2. | Does the instructor seem interested in the class? | Very much | 1 | 2 | 3 | 4 | 5 | Not at all |
| 3. | Do you feel that the instructor is aware of you as an individual? | Very much | 1 | 2 | 3 | 4 | 5 | Not at all |
| 4. | Are assignments clearly given? | Always | 1 | 2 | 3 | 4 | 5 | Never |
| 5. | Does the instructor make clear which concepts are most important? | Always | 1 | 2 | 3 | 4 | 5 | Never |
| 6. | Does the instructor spend time wisely, staying on track and emphasizing the important points? | Always | 1 | 2 | 3 | 4 | 5 | Never |
| 7. | The instructor's explanations are | Clear and complete | 1 | 2 | 3 | 4 | 5 | Vague and incomplete |
| 8. | Does the instructor use the board effectively? | Always | 1 | 2 | 3 | 4 | 5 | Never |
| 9. | Does the instructor generally encourage class participation? | Always | 1 | 2 | 3 | 4 | 5 | Never |
| 10. | Is the instructor in control of the class? | Always | 1 | 2 | 3 | 4 | 5 | Never |
| 11. | Does the instructor speak so that you can hear and understand him/her? | Always | 1 | 2 | 3 | 4 | 5 | Never |

12. Are the instructor's quizzes
 too hard? Always 1 2 3 4 5 Never
 too easy? Always 1 2 3 4 5 Never
 ambiguous? Always 1 2 3 4 5 Never
 helpful in learning the
 material? Always 1 2 3 4 5 Never
13. Have you observed cheating Yes No
 in your class?
14. What could the instructor do to help you learn more effectively?
15. Overall rating Excellent 1 2 3 4 5 Poor
16. What grade do you expect to A B C D E
 achieve in this class?

Additional Comments:

PART V: DISCUSSION OF ACTIVE LEARNING QUESTIONS

The Active Learning questions are designed to be considered by groups of students in class. Often they work well for introducing a particular topic in class. The purpose of these questions is to give the students the opportunity to discuss their ideas. They also give us as teachers the opportunity to hear these ideas. This is not done solely to get the students to a correct answer for a particular problem, but rather to try to get a better feel for the students' conceptual understandings of the topics we teach in chemistry. Some of these problems have definite answers and some do not. In the discussion in Part V, we will provide you with rationale for the questions, although not always clear-cut answers.

The answers to these questions are not in the *Complete Solutions Guide* for a reason. Simply stated, students tend to look up answers much too quickly, and the point of these questions is not to merely achieve an answer. Therefore, we stress that you emphasize the students' ideas more than just the correct answer; telling the students the answers to these questions is an inefficient way of trying to get them to understand. All of these questions have been used with students at UIUC, and many of the multiple-choice responses have been developed from open-ended questions answered by UIUC students. In addition, many of the multiple-choice questions have been tried with several hundred entering freshmen, and the results will be given in the discussion of the particular questions.

CHAPTER ONE

1.
 - a) 43,800 minutes
 - b) 40,320 minutes
 - c) These are different due to the differences in conversion factors. Answer "a" could be said to be more "correct" because the conversion factor "4 weeks/1 month" in part "b" is less exact than those in response "a." The problem is useful to show the students that they can get different answers with different conversion factors, and that these conversion factors have an infinite number of significant figures.
2. The problem is useful at this point to stress the idea of dimensional analysis and to show the students that they have an intuitive way of dealing with these types of problems, especially if they have a strong understanding of the problem (which almost all of the students will in this case). When you get to Chapter 3, you can go back to this problem and make it an analogy (although a crude one) with chemical stoichiometry; the idea of a limiting "reactant" is also introduced (you can even write out an equation showing $3P + 2D \rightarrow 4C$). The answers are a) 210 pennies, b) 140 dimes, c) 321 g, d) 280 pieces of candy, e) 2.86×10^3 g, f) 280 pieces of candy.

3. Multiple-choice questions such as this one will be used occasionally, and have proven to be good questions in that the students find explaining away the wrong choices problematic more so than choosing the correct one. Students at this level are generally rather “test savvy” and are quite adept at choosing the right answer; they can even justify it to some extent. However, these choices are the most popular actual student responses to open-ended questions given by chemistry students at UIUC.

Choice “c” is the best response, although it does not explain why an object that weighs more than an equivalent volume of water would sink. One of the benefits of these types of problems is to let the students know that there can always be a deeper level of an answer. Choice “a” is a very popular answer, presumably because many students have an intuitive model that an object floats due to being suspended by surface tension (many students will even use the formula $P = F/A$, or discuss the use of snowshoes in justifying this answer); this also explains choice “d” to an extent. Choices “b” and “e” are chosen by students either because of a confusion with mass and density, or because the equation for density includes mass and volume. More than stressing the answer, however, is finding what the students think about density and how they use it as an explanatory device. Merely telling them the answer will mean little if they have not had time to think about it and discuss it.

4. For part a, choice “iii” is the correct choice, in concurrence with the law of conservation of mass. Many students (almost 30% of several hundred incoming freshman tested at UIUC) chose “iv” or “v”, indicating that the mass decreased. Some felt that this was due to a “reaction” of some sort, while many believed that in any transfer one would “lose” some of the solution.

For part b, “v” is the best answer (although “iv” is reasonable too, especially since the term “somewhat” is rather vague). Almost 40% of the several hundred incoming UIUC students given this problem believed that the volume would remain the same. This introduces the idea that there are “spaces” between molecules.

5. Of several hundred incoming UIUC students, the breakdown for responses for part a was as follows: i. 7%, ii. 40%, iii. 12%, iv. 37%, v. 3%.

Note that the most popular response is that the bubbles are hydrogen and oxygen gases, which makes a response of “chemical change” to part b more reasonable. It seems, then, that if we merely tell the students that the boiling of water is a chemical change, yet they believe that the product is hydrogen and oxygen gases, they will be more confused than they were initially. One of the indicators of a chemical change (with which the students may be familiar) is that a gas is given off, which, from their perspective, is what occurs as water boils.

6. If possible, it might be a good idea to either demonstrate this reaction to the class or have groups of students observe it. Again, students may have good reasons for listing a process as a chemical change or physical change even if they are incorrect. In many ways, their reasons for their answers are more telling than the answers themselves.

In this case, the wax is undergoing both chemical and physical changes as it is both melting and burning. The wick is burning, and is therefore undergoing a chemical change (although it also looks physically different). The glass rod is not really being changed significantly, although it is becoming coated with carbon.

7. This question helps the students to see the limitations of any classification system while at the same time allowing you to get a better idea of how they are using the system. However they classify each substance, it is more important to know which characteristics of a solid, a liquid, and a gas they are using. For example, a bucketful of sand takes up the shape of its container, and students may even claim that pudding is relatively compressible. After allowing the students sufficient time to work on this question in groups, you may wish to ask them about toothpaste.
8. Many times students who know the rules of significant figures have difficulty with them in practice and do not realize what it means to limit the precision of a measurement. The answer will be dependent somewhat on how the students read the volume in each of the graduated cylinders. What is important is that they realize that they can be as precise as one-tenth of a milliliter in the left cylinder and to one-hundredth of a milliliter in the right cylinder. Thus, if the student says that the volume of the left cylinder is 3.0 mL and the volume of the right cylinder is 0.31 mL, the total is 3.3 mL. You may wish to have a few more of these problems available for the students to try, and be prepared for the student to be a bit upset because, in their minds " $3 + .31 = 3.3$," for example. While many of them have intellectualized the rules, actually dealing with a substance such as water will frustrate some students.
9. Yes, Paracelsus is admonishing his students to learn by observation. Observation of facts is the foundation for all of the other steps in the scientific method.
10. Experimental results are the facts with which we deal. Theories are our attempt to rationalize those facts. If the experiment is done properly and the theory can't account for the facts, then the theory is wrong. The idea of a model or theory is a major emphasis of this text, and this idea that "theories are truth" seems to be strongly ingrained in students.

11. If a measurement is not precise, then it may be close to a particular “true” value only by chance. The next measurement may be way off or even the next several may be far from the true value.
12. Problems such as these are good in that they require the students to decide what information is necessary and how to use this information. Students can often enter numbers into an equation correctly if we give them only the necessary numbers and the equation. However, a problem such as this one requires the student to think about the problem-solving approach that is needed.
13. Students are often confused about significant figures, and many times even those students who can use them correctly are not sure about the purpose or application. Problems such as these require that the student understand, not merely know, the rules of significant figures, especially with the application to lab measurements.
14. Mass and volume are often confused or used interchangeably. While both objects will displace the same amount of water, many students will think the lead will displace more water since it is “heavier.” This problem also allows for a discussion of measuring volume by water displacement.
15. This is a good time to have the students draw molecular-level pictures. The drawings that the students submit will go much further in letting us know their understanding of the concept than their reciting a text definition.
16. You travel almost 100 feet (about 95 feet in a second). There are 60 seconds in a minute; 60 minutes in an hour; you travel 65 miles in an hour; and there are 5280 feet in a mile.
17. A mathematician would say “.43.4” and, if these are measurements, a scientist would say “43.” If, for example, we add 15.4 mL of water to 28 mL of water, we are saying 15.4 ± 0.1 and 28 ± 1 .
18. A mathematician would say “.430.466” and, if these are measurements, a scientist would say “430.” We can only have as many significant figures in the answer as the least number in one of the measurements when we are multiplying numbers. There is uncertainty inherent in the measurements.

CHAPTER TWO

1. Over 70% of the several hundred incoming UIUC students correctly chose response e. However, almost 25% of the students chose response “d.” Again, the largest benefits of a question such as this one are to have the students explain away what they believe to be incorrect explanations and to hear the students' ideas. For example, by choosing d we could believe that the students

do not understand an atomic theory. On the other hand, they may simply be confusing the terms “atom” and “element.” These types of insights are important for us to make.

2. This question should get the students to start thinking about the conversion between mass and number (which will be useful for the mole concept). One way of solving this problem would be to find the mass of the chalk before and after writing one's name, and find, by subtraction, the mass of chalk it takes to write the name. Then, by knowing the mass of a “chalk molecule,” one could calculate the number of “chalk molecules.”
3. The main purpose of this question is to get the students to appreciate, in general, the purposes of a model and to consider, more specifically, Thomson's model. The answers to this question will vary from student to student, and it is important here is to stress that the answers are more contingent on their support than on absolute “correctness.” However, most students will probably claim in part a that electrons are the most important for the formation of compounds (especially due to Thomson's findings that the electrons are so easily removed). In the case of part c, it is most important to stress that any model that explains Thomson's findings is reasonable, and that this endeavor is what developing models is about; that is, matching observations, not reality. This emphasis on models is extremely important in chemistry and will be further stressed in Chapter 5.
4. When several hundred incoming UIUC students were asked what happens to the size of molecules and to the mass of a sample as ice is heated (until it is steam), about one-third of the students believed that the mass and molecular size varies. Most of these students believed that the mass of the ice is greater than that of the steam, and that the size of the “steam molecules” is greater than that of the “ice molecules.” The drawings in conjunction with the explanations can reveal quite a bit about the students' ideas.
5. When several hundred incoming UIUC students were asked this question, almost 40% of the students responded that the mass of this system would decrease, presumably because the chemical has either “disappeared” or has been converted into a gas (which many believed has less mass than the solid). Most of the rest correctly believe that the system ends up with the same mass, although it again is interesting to ask them explain why (especially without using the phrase “law of conservation of mass”).
6. Answer “b” is correct. Some students will believe that the 2 in the formula suggests something about the mass of hydrogen or oxygen. While this is not widespread (especially of chemistry majors) it should be addressed if it comes up.
7. There should be no difference. It does not matter how a substance is produced,

it is still that substance.

8. Yes, many questions are raised from Dalton's theory. For example: What are the masses of the atoms? Are atoms really structureless? What forces hold atoms together in compounds?
9. We now know that some atoms of the same element have different masses. We have had to include the existence of isotopes in our models.
- 10-13. Students all too often memorize rules for naming compounds without understanding that the names are systematic and that the rules are made to be as simple as possible. For example, Roman numerals and prefixes are used only when necessary (that is, if there is more than one possible compound to fit the name, as with "iron oxide" for example). The more we stress the purpose of rules with students, the more they expect there to be a purpose to these rules, and the greater the chance they will achieve understanding.
14. Students should realize that it is the number of protons that "define" an element. Thus, choice "d" is correct (because choices b and c are correct).
15. Some students will believe that "a" and "c" are correct, although only "c" is correct. While this is not widespread (especially of chemistry majors) it should be addressed if it comes up.

CHAPTER THREE

1. As stated in the question, the given choices are actual student responses (incoming students to UIUC). Again, the most important goal of a problem such as this is to get the students talking about why they believe certain responses to be incorrect. Most students will undoubtedly know that d is correct, but many of them will also be surprised that some of the other responses (especially a and e it seems) are not correct.
2. The main purpose of this question is to have the students come to the same definitions of what is meant by a formula and what is meant by an equation (many times students will use these terms interchangeably). Also, if the students are more clear about what information is given by a formula (the elements in the compound and the ratio of different atoms in the molecule), they will be more likely to understand what is meant by a balanced equation.
3. This can be thought of as the next step in the chemical stoichiometry analogy (the first step being question 2 from Chapter 1). Many students will be able to do this question correctly, even regarding the limiting reactant portion, yet still claim not to understand other problems (chemical). It is therefore a good idea to make the analogy explicit once the students are successful at this question (which they

often are).

4. Even when the students can balance equations and perform limiting reactant calculations, many times they will have difficulty drawing representations of the reactions. In fact, when incoming students at UIUC were given an open-ended question asking them to represent the equation in this question, many of them drew “tri-atomic” hydrogen (presumably because there is a coefficient of three with respect to the hydrogen). In the multiple-choice version given to several hundred students, almost 25% of them chose answers with tri-atomic hydrogen, and only slightly more than half chose the correct response. Their solution should include 4 molecules of ammonia and 4 molecules of diatomic nitrogen.
5. Research on several hundred incoming UIUC students suggests that many of the students will choose “a” from the correct drawing in question number 4 (almost two-thirds chose it). While it simplifies to the correct answer “d” it becomes problematic if students try to use this equation for stoichiometric calculations (there is a 1:1 mol ratio between N_2 and H_2 , for example). Choices “d” and “e” are both acceptable (as is “b,” although it is not balanced). This question serves as a good springboard for a discussion of what a chemical equation represents (for example, the excess reactant is not to be shown on the right side of the equation).
6. This type of problem is good in that it makes the students really think about what they need to know to solve a problem, as opposed to trying to figure out how to “plug and chug” the numbers to arrive at a correct answer. In this case, they need to know the balanced equation, the molar masses of A and B, and the molar mass of the product that is formed. Have the students explain both how they will use the information and why they need the information they claim to need.
7. In this case, only about 10% of the several hundred UIUC students given a multiple-choice version of this type of question correctly chose a response in which the gases had a mass greater than that of the object being burned (because it is becoming an oxide—or oxides—of that sample). In fact, about half of the students believed that the mass of the gases would weigh the same as the sample that was being burned (presumably due to the memorized law of conservation of mass).
8. Results from several hundred incoming students at UIUC reveal that about half of the students believed that the mass of an iron bar decreased when it rusts, and almost 25% believed that the mass stayed the same. Less than 20% correctly chose “b.” Again, it is important in a question like this one to allow the students to discuss their ideas, especially in deciding which choices they believe are incorrect and why.

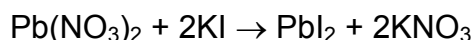
9. For this question, over one-third of several hundred incoming UIUC students incorrectly believed that the answer to this is “yes.” More important in this case, however, is to get the students to verbalize their explanation.
- 10, 11. These questions are related to question 4 from Chapter 1. Only about half of the several hundred incoming UIUC students correctly chose “d” for question 10. Answer “d” is the best choice for number 11 (although “e” is a reasonable choice as well). About half of the students chose “d”, and 20% chose “e.” Again, it is important to get at the reasons for the answers, especially since students seem stuck on the law of conservation of mass, although not necessarily in this case. Also, in the case of question 11, over 10% chose “a,” presumably because more of chemical A (by mass) is involved in the reaction (only about 3% chose “b”). As with many of the other questions, it is a good idea to use this question to better understand the students' ideas.
12. Many students believe that because the formula for water is H_2O , a 2:1 mixture of hydrogen and oxygen gases is equivalent to water. Having the students draw molecular-level pictures of the mixture and the compound should help students to relate the macroscopic to the microscopic, which is an important idea in chemistry. This question also gets to the differences between a mixture and a compound.
13. The periodic table shows the average mass of all of the isotopes. Therefore ^{35}Cl must be more abundant because the average mass is closer to 35 than 37. You may also wish to have the students calculate the relative abundance of each of these two isotopes.
14. This question could be followed by a discussion of what average mass means. In this case, a is correct—no carbon atom has a mass of 12.011.
15. Even students who can balance equations often have difficulty with these questions. The student has to understand why we balance equations, and what the coefficients and subscripts symbolize before they can answer them. After all, a balanced equation is just letters and numbers with an arrow. Its meaning is not intuitively obvious.
16. Many students say 1.0 or 2.0 moles. Some students are uncomfortable with the answer 0.50 mol, presumably because it is a fraction.
17. Students often use the law of conservation of mass inappropriately. It is sometimes a good idea to give the students a question that seems to contradict a known truth. In defending their answer, students come to a deeper understanding of the concepts.
18. The answers are “a” and “c.”

19. While many students at this level can solve limiting reactant problems, many can do so without an understanding but a mere algorithm. This requires the students to understand what is meant by a limiting reactant. In this case, two molecules of (the second reactant) are left over.
20. Although many students at this level have had a great deal of chemistry instruction, a lot of the students do not understand why we use the mole concept in chemistry. Many students cannot explain what a mole is or what it is used for. In answering this question, students will often claim the mole is "how much the compound weighs," or something like "a mole is the mass of the atoms in a compound." Students do not always appreciate that the mole is used to equate mass and number so that we can, in effect, count by weighing.
21. The answer is "c." Limiting reactant problems can be difficult for students, although many students at this level can solve them. Students who have difficulty with these often try to memorize an algorithm and may multiply by an incorrect ratio of coefficients. This type of question requires that the students think about which reactant is limiting and why. It is important to have the students explain their answers.
22. Students who have difficulty with limiting reactant problems often confuse the amount of a reactant they have with the amount that is needed for a complete reaction. In this problem, the answer is "b."

CHAPTER FOUR

1. This question is consistent with the many sections that emphasize pictorial representations of ions, molecules, and reactions in solutions, and stems from research that points to the fact that while many students are successful at balancing equations, they are not able to represent these equations. The main point here is to use the students' drawings to get a better idea into their thoughts.
2. When several hundred UI students were given this question, almost 60% correctly stated that the concentration increased. The remaining 40% were about evenly split between claiming the concentration either decreased or stayed the same. This belief presumably came about because the students are not visualizing what is occurring when the water in a solution dissolves and/or because they are using an incorrect definition of concentration.
3. This question is intended to get the students thinking qualitatively (and perhaps semi-quantitatively) about concentration. Again, students are generally good problem solvers in an algorithmic sense, although they do not always consider the concepts in the problems. In this case, there is 3 times as much sugar in solution A as in solution B; the volume of solution A is 1.5 times as great as that of solution B; and the concentration of solution A is 2 times that of solution B.

4. This problem is intended to get the students to consider a representation of a reaction in solution. Given the chemical equation, most of your students can undoubtedly balance this equation, although it is not clear how many students will know what to do when you ask them to draw a representation. The equation is



5. This question allows the student to see that the same element can have many different oxidation states, even without being an ion.
6. Students often want mnemonics (and use some with redox questions). While these may help the student know the “what,” they do little, if anything for the “why.” By gaining electrons, the oxidation state is reduced, because electrons are negatively charged. Thus “gaining” and “reduced” make sense in this context, and one need not rely on a mnemonic device.
7. This question requires the students to understand what the mole-to-mole ratio means. Often students will incorrectly use a coefficient (multiply by the coefficient instead of divide, for example). In this case, the answer is “b.”
8. This question is another example of having the students use pictures to convey their ideas. These should be accompanied, as always, by an explanation. Look for a confusion of concentration and volume or concentration and amount of solute. The fact that concentration is a ratio makes it confusing for many students.
9. In this case, the person with the solid should add 0.88 g of NaCl to 150.0 mL solution. The lab partner should add 6.0 mL of 2.5 M NaCl to 144 mL of water.
10. This problem requires the students to think about the application of the solubility rules. The answer to this is “a.” The precipitate formed is barium sulfate, and the other product is water. Thus, the ion concentration decreases, and the bulb grows dim.
11. Students find problems such as this (in which students must decide which data are needed to answer a question) difficult. The molar mass of A is not needed because the students do not need to solve for the concentration, they just need to compare relative concentrations. Choices “a” and “c” will suffice, although “d” would need to be known for an actual calculation.
12. Similar to question 7, this problem requires the students to decide which data are needed. Choices “a” and “d” need to be known.

CHAPTER FIVE

1. In this question, the membrane should look concave (pushed into the test tube) in both cases. Even students who believe that the membrane is pushed into the test tube initially sometimes claim that the membrane will “bounce back” unless the syringe is continually pushed down. A good variation with this problem is to use a similar set-up but with the syringe coming from the side, as many students believe that pressure is unidirectional (so that pushing in a syringe from the side will not cause the membrane to be pushed inward).
2. When several hundred incoming students at UIUC were asked this question the breakdown (percentages) for the answers was as follows: 65, 20, 10, 3, 2. Note that only 10% of the students correctly answered this question, although all of these students had at least one year of high school chemistry. Most students believe that there is air inside the tube, and both responses “a” and “b” imply that the weight of the mercury is inconsequential (because these responses have the pressures equalizing). Again, the point is not to simply tell the students the correct answer (most of them have undoubtedly heard it) but to allow them to discuss the reasons for their answers so that you can get a better idea of their models. As with other questions, it seems that the most important consideration is to have the students try to explain why they believe that the responses they do not choose are incorrect; they are good at supporting answers that they believe to be correct but find it difficult to explain away other responses. It has proven to be enlightening to ask the students how they would make a barometer (especially with reference to the amount of mercury that is added to the tube).
3. Interestingly, although most students will choose “a” or “b” in the previous question, they will, when focused on the amount of mercury in the barometer, claim that the weight does matter (many will have no problem seeing a difference between a mercury barometer and a water barometer). However, they usually have difficulty grasping the idea of pressure as being measured per an area, and, for example, many students will claim that the level of mercury in a tube with a greater circumference will be lower than that in a tube with a smaller circumference. In fact, the level of mercury will be the same in all of the barometers.
4. In this case, the density of a gas will not change if the sample is kept at constant volume, but it will if it is placed in a container fitted with a piston and the temperature is increased (in that case, the density will decrease because an increase in temperature increases the volume). Many students believe that the density of a substance cannot be changed (even after doing calculations of gas density). It would be a good idea to have the students draw pictures of these situations, as emphasized in the chapter.
5. Of several hundred incoming UIUC students given this question, just over half correctly claimed that nothing is between the “dots.” However, almost a third of

the students believed that there was either air or oxygen between the air molecules. Again, this shows the need for having the students consider magnified representations of gases (as emphasized in the chapter and in these questions) because while most students can solve problems in which they are to plug numbers into the equation $PV=nRT$, almost a third of the students believe that air molecules are located in the space between air molecules.

6. Chemistry majors at UIUC have been asked to discuss this question, and initially most claim that they have done this with straws for years; however, they are hard pressed to explain it, especially when asked to describe this phenomenon by using the kinetic molecular theory. Actual student models that have been proposed include one's finger blocking gravity; and because there is less air in the straw than in the room, there is less pressure in the straw, and the water is pushed up (at the surface they are correct by claiming that there is less pressure, their explanation is incorrect). It might be a good idea to have the materials available (straws and water) so that the students can observe this phenomenon; many notice that a drop (or so) of water comes out of the straw, and some are able to then make the connection that because of this increase in volume of air in the straw (with the same amount of air) the pressure decreases. However, the point is not about telling the students the right answer, but allowing the students to discuss their ideas.
7. Many students are confused by this question and some almost believe that this is proof that Boyle's law doesn't work (in fact, one student claimed that this is true because air is a real gas, and Boyle's law is for ideal gases). It seems that many student use either Boyle's or Charles's law without consideration of the conditions (in this case because more air is added, both pressure and volume can increase, as long as the tire is pliable).
8. It might be a good idea to have the students draw a picture of this situation. By this point they should see the advantages of pictures on their own, but you may have to remind them of it. In this problem, two moles of gas become one mole, and according to the KMT, this gas (XY) should take up half the volume of X and Y separately. Of several hundred incoming UIUC students given this question, almost half (47%) claimed that the volume would not change, and the remaining half were evenly split between the volume increasing and the volume decreasing.
9. The most popular response of several hundred incoming UIUC students to this question is "b" (about 40%)—that hot air rises. The correct response "d" was chosen by only 20% of the students, and some of them chose it merely because it included the term "density" and they knew this should be part of the answer. Choice "e" was also popular (23%). Again, the emphasis should be on letting the students discuss their ideas and having them explain why they believe the choices they did not choose are incorrect. Merely telling the students that the answer is "d" will not help the students understand the nature of gases.

10. When several hundred incoming UIUC students were asked this question, only about 27% correctly claimed that the situations would all look the same with a static drawing by merely changing the temperature. The most popular response (42%) was that when cooling the gas the molecules will become significantly closer together. Again, the emphasis is to have the students represent gas molecules and the KMT. In the case of evacuating the gas, most students will draw a picture with fewer gas particles evenly spread (not all pay attention to the detail of the actual number of “dots”), although some leave a lot of open space at the top of the drawing (which makes sense if they think this empty space is air anyway, as cited in question 5).
11. In this case pressure outside the balloon is decreasing, and thus the balloon will expand to equalize pressures. It may be a good idea to compare this question to the barometer question (number 3) in that the students see these situations as similar—merely an equalizing of pressures.
12. The question is essentially having them explain the ideal gas law, so do not let them use $PV = nRT$ to explain it. That is, have them explain why $PV = nRT$ works. The goal is to help them understand the models that we use when dealing with gases.
13. In this question, the main problem is that energy and velocity are not the same. That is, the velocity depends on the mass of the particle as well.
14. Students will often relate pressure to volume or pressure to the amount of gas present without considering the other variables. Once the students have considered this question, ask them to discuss the same situation when the balloon is larger or smaller than the flask.
15. For Dalton’s law of partial pressures to hold true, the simplifications of no interactions among gas particles and negligible particle size must hold true. Have the students explain why.
16. Students often will memorize the equation relating the density of a gas and its molar mass without considering from where it comes. Do not allow the students merely to use this equation as an explanation.
- 17,18. These problems require the students to think about the equations rather than merely plug numbers into them. Many students find questions like this difficult, but it is a good frustration for them if you require them to explain their answers without complete equations. For example, to compare kinetic energies, we only need to know temperature. The number of collisions and force per collision also require we know the mass of the molecule (and the particle size for collisions with other particles).

19. Look for more spacing between particles in the gas than in the liquid, and the least between particles in a solid. Given the same substance, the particle sizes should be the same.

CHAPTER SIX

1. Students tend to like questions such as these, which discuss something familiar to them and which can be explained with what they have learned. The key to this question is the differing heat capacities of the different substances. It is important that the students are not allowed to merely say, “because of the heat capacities,” but must explain their ideas.
2. Students have difficulty with this question, and for good reason—consider that if compound A is lower in energy than compound B, it could be said that the bonds in A are stronger than those in B; thus it takes more energy to break apart something that is lower in energy. This idea is contradictory to some students. The given example (with the hydrogen, oxygen, and water) is one that can be referred to over the course of the semester whenever the phrase “lower in energy” comes up. We have found it helpful to have the students observe this reaction so that it is very clear that energy is given off as the reaction proceeds. They can see that the water, because it is formed naturally, is, by definition, lower in energy (and more stable) than the hydrogen and oxygen.
3. The main purpose of this question is to see how the students use these terms in the discussion. Quite a few correct answers are possible, but consistency of terms should be stressed.
4. If you ask students whether you make ice by putting water in an oven or a freezer they will always get it right (unless they are being sarcastic in response to the ridiculousness of the question). Ask them if the freezing of water is endothermic or exothermic and why, however, and you may be surprised by their responses. Many students initially (and mistakenly) believe this process to be endothermic, and it appears this is the case because ice is cold. Exothermicity is often equated with things blowing up and lots of heat given off, so the students have difficulty accepting this process as an exothermic process. (This, of course, assumes that we take the water to be the system, which is why it is important to listen to the students as they define these terms in context.) As with other questions, merely giving them this answer will probably not suffice in increasing their understanding—they need to have time and discussion to work through their confusion.
5. Because we assume the second statement to be the case when doing calculations with calorimeters, many students believe that heat is conserved. This is problematic because assuming heat is conserved reinforces the mistaken idea that heat is a substance.

6. It may be a good idea to have the students work together on a problem such as this one when it is first introduced. If you decide to have the students work on this question, it is strongly recommended that you have them discuss it in question 7.
 - a) $w = -7.9 \text{ l. atm}$
 - b) $w = -9.0 \text{ l. atm}$
 - c) $w = -9.1 \text{ l. atm}$
7. Work is not a state function simply because its value is dependent on the pathway. In a one-step expansion, all of the work is done against a relatively small external pressure; in two steps, part of the work is done against a larger pressure. Have the students experiment with their own numbers and they may find that the magnitude of work is maximized in a two-step process when the volume of the first step is exactly in the middle of the initial and final volumes. Even if they do not figure this out, you can tell them and have them try to explain it, which would be a good opportunity to have them make PV diagrams.
8. Water acts as a greenhouse gas. Because of this, the temperature fluctuations of oceans are small as compared to temperature fluctuations of air. Thus oceans can act as a heat reservoir for areas close to them. This is another example of a question that has the students consider a “real-world” situation.
9. If Hess's law were not true it would be possible to create energy by simple reversing a reaction using a different series of steps. This question helps the students to see the interconnectedness of much of what they learn in chemistry; this is to oppose many students' perceptions that chemistry consists of a group of disparate facts.
10. The negative sign is a convention and the students should realize this. If the volume of an ideal gas increase, work is done on the surroundings, thus is given a negative sign.

CHAPTER SEVEN

1. The purpose of this question is to prompt students to think about the difference between wavelike and particulate properties. It is important for the students to realize that the properties describe the same electron, and perhaps even to realize that the concept of an electron is a model we use.
2. This question is primarily to get the students thinking about what it means for something to be a model, why a model is said to work, and why one would want to abandon an existing model in favor of a new model. In this case, Bohr's model works with hydrogen (and with any one-electron species). In the one-electron cases, it is quite a good model that can predict behavior (although models do not

describe reality). Because of the complexities of electron-electron repulsion, however, the model fails to account for other elements.

3. In this problem the students should note that there is a great change between the second and third ionization energies for both elements X and Y. This could indicate that the elements are alkaline earth metals. Because the ionization energies of Y are greater than those of X, element Y should be above X in the periodic table (so, for example, Y could be Mg and X could be Ca). The purpose here is to have the students think about what these numbers mean and how to use the learned trends.
4. Many students think of the first ionization energy for a noble gas as being higher than the second ionization energy because of a memorized rule such as “All elements want to be like noble gases.” While this statement is true to an extent, many students misuse it to mean that an atom such as He “really does not want to” lose an electron; however, if it already has, then it is not so difficult to lose the next one. Using this logic, students will also claim that the second ionization energy for an atom such as Mg will actually be negative so that it can be like Ne. Questions 5 and 11 in this chapter and question 5 in the next chapter also get at this idea. Again, the goal here is to prompt the students to elaborate on their ideas.
5. Many students realize that taking a second electron from lithium will be difficult (take quite a bit of energy) because after losing the first electron, the ion will have the same configuration as He. However, as mentioned above, many students will also claim that the second ionization energy for beryllium will be negative (exothermic) because the beryllium ion will “want” to lose another electron to be like He. Do not allow the students to look up the numbers before answering this question (although it might be a good idea to do so after a lengthy discussion) and make sure that the students do not just provide a quick answer (such as “The second IE for Li is greater”) but rather explain their reasoning.
6. This question is meant to stress exceptions. Even though these exceptions are explained at length in the book, many time students tend not to give graphs and figures their full attention; whenever you can ask them to explain a figure or graph, it is probably a good idea. This will be especially true in later chapters when some Active Learning Questions explicitly ask students to explain figures.
7. This question is intended to get the students to think about how to use the trends they have learned. Also, you can get a better idea about what they are thinking in terms of atoms “wanting” to be like noble gases as in questions 4 and 5. Having the students compare their answers to those in the book and then trying to understand the differences can also be enlightening.

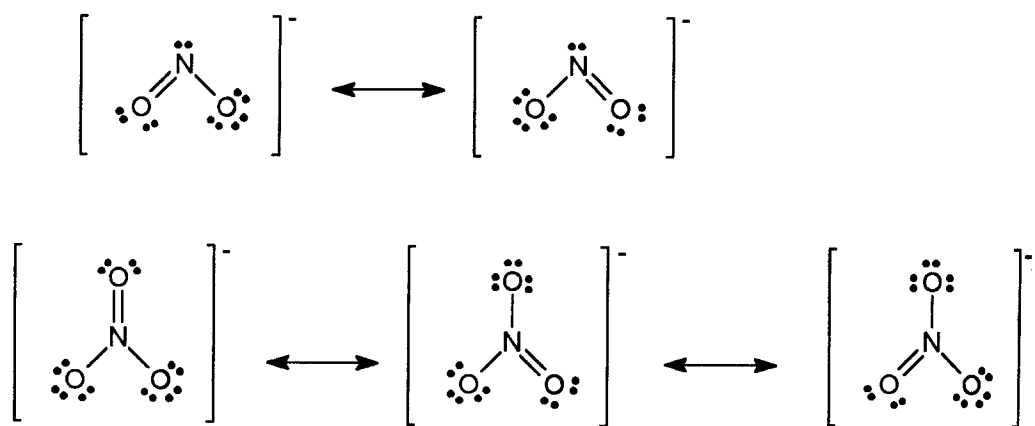
8. Students have noticed this trend and are not sure why it exists. In fact, some students have just accepted it as “the way that it is,” which is an attitude that should be dispelled as soon as possible. Questions such as this one, which ask the students to explain something that they may have been merely accepting, are therefore quite educational. Even if we are not sure of the answer, it is a good idea for the students to see that we can make educated hypotheses based on previous information. Students may equate metallic behavior with atomic size and answer accordingly, or you may have to lead them through this concept a bit; whatever the case, allow the students time to discuss their ideas.
9. This question is intended to get the students thinking about the vagueness inherent in the concept of an electron, away from a negatively charged particle with a known orbit.
10. Again, one of the best aspects of a question like this one is that the students are required to explain why they did not choose four of the choices, and, of course, why they did choose a specific answer (not just to choose it but to explain it). In this case, the answer is “c,” but what is more important at this point is to allow the students the chance to elaborate on their ideas. Once a considerable discussion has taken place, and after choice “c” is considered as the best answer, this situation serves as another example of Hess' law (depending on when it is covered).
11. As with other questions for this chapter, the goal here is to get to students' ideas about ionization energy when an atom losing an electron is isoelectronic with a noble gas. The most important issue here (as with the other questions) is not to just allow students to answer either “true” or “false” but to explain their ideas. In this case, energy is still required for a potassium atom to lose an electron even though many students will claim that a potassium atom “wants to do this,” and will therefore lose an electron spontaneously (sometimes even claiming that energy is given off).
12. The main reason for this question stems from the observation that many times students merely memorize trends across and down the periodic table or, at best, use a memorized rule as an explanation (such as “there is an increase in electrons”). Thus this question provides a discrepant situation—if a student uses the “increase in electron” rule as an explanation, this question serves to show the students that this rule has the opposite effects intended. Have the students explain the trends instead of just memorizing them.
13. Students have a notion that we know exactly where the electrons are and that they move in predictable orbits. We need to help them accept the idea that our models are based on probability.

14. The term “orbital” is unfortunate in that many students infer a circular orbit from it. However, it represents a region of probability in which one is likely to find an electron. This distinction is crucial.
15. These concepts are discussed in section 7.5, but the subtleties can be missed without an “active” read of this section.
16. This is true although some students will say it is false and claim that the hydrogen atom only has a 1s orbital. This question can be followed by a discussion of the concept of orbitals.
17. This 2s and 2p orbitals are degenerate for hydrogen because there is only one electron. For helium, the 2s orbital is lower in energy than the 2p. If one electron of the helium atom is excited, it has a greater probability of being in the 2s orbital because of the so-called “penetration effect” discussed in section 7.9.
20. The students should be able to do this with data given in Chapter 7, providing they understand the concepts of ionization energy and electron affinity. The data needed are: ionization energy of Na (495 kJ/mol), ionization energy of F (about 1750 kJ/mol), and electron affinity of F (-327.8 kJ/mol). Thus, to take an electron from Na, and then have the F take it in, would require 167.2 kJ/mol (495-327.8), while taking from one F and giving to another would require about 1422.2 kJ/mol (1750-327.8). While both are energetically unfavorable, it is “less unfavorable” for the reaction between Na and F than F and F. This is a good lead into a discussion of lattice energy for Chapter 8.

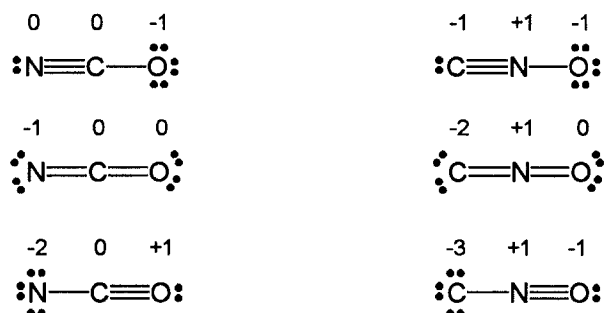
CHAPTER EIGHT

1. It would be best to have the students answer this question only using a periodic table (without the textbook). Having them understand how these properties are related will decrease the likelihood that students merely memorize the trends.
2. This question allows the students to discuss lattice energy and ionization energy, and to try to understand how to achieve a lowest possible energy state. A discussion of this question may also shed some light on how an alkali metal can “want” to lose an electron, even though the IE is endothermic. Students may be able to do quantitative problems with lattice energy, but they are not always sure what lattice energy means; this problem is meant to get them to apply their definition.
3. As with question 1, have the students answer this question without using the textbook. The goal here is for students to explain their reasoning rather than simply use a catch-phrase. Allow the students to discuss their predictions once they have looked up the answers in the text, as this discussion could provide you with more insight into their ideas.

4. To answer this question, the students should draw Lewis structures of these molecules and note that while the shapes are the same, the atoms (Br versus H) are different. The C-H bond energy is therefore dependent on the environment. The answer reflects the larger electronegativity difference between Br and C than between H and C (the polarity of the bonds makes the C-H bond weaker in CHBr_3). Do not allow the students merely to claim that the answer has to do with electronegativity differences, but have them explain their reasoning.
5. As mentioned in the last chapter, students sometimes believe that the second IE for Mg (for example) will be exothermic because, by losing a second electron, the Mg atom becomes isoelectronic with a noble gas. In this case, many students, consistent with their ideas about the second IEs, will claim that the answer to this question is true, which is not the case. By losing an electron, the positively charged oxygen ion becomes less likely to lose another electron.
6. In drawing the resonance structures, the NO_2^- can be thought of as having an average of $3/2$ bonds, and the NO_3^- can be thought of as having $4/3$ bonds. Thus, the NO_3^- has a longer bond length.



7. Drawing the possible structures for these ions, we get (with formal charges):



Given formal charges, the first structure is the best representation for each.

8. No, we would expect the more highly charged ions to have a greater attraction for electrons. The electronegativity of an element depends on its oxidation state.
9. The electron-electron repulsions are much greater than the attraction of the electron for the nucleus.
10. Instead of just knowing the “what,” the students should also consider the “why?” Many times students will memorize that we sometimes represent a chemical bond as “two shared electrons” but do not consider why this bond forms. A discussion of this issue could help to introduce or reinforce concepts in thermodynamics.
11. Students often memorize the definitions of “ionic” and “covalent” but cannot explain the difference between them. Again, asking the students “why?” instead of “what?” gets to a deeper level of understanding.
12. A bond between Na and Cl is considered to be ionic (an electron is transferred from the Na to the Cl and electrostatic forces hold the ions together). With C and O, electrons are shared, but unequally (a polar covalent bond) in which the electron density is greater around the oxygen. For N and N, electrons are shared equally (a covalent bond).
13. The students need to understand the difference between the polarity of bonds and the overall polarity of the molecule. Thus, CH₄ and CCl₄ are equally non-polar as molecules, although CCl₄ has polar bonds. CF₂H₂ is the most polar, followed closely by CCl₂H₂. CF₂Cl₂ is slightly polar.
14. Many first-time students think the answer to this question is “yes” and can “prove” it with simple examples such as water. However, a discussion of this question can get students to see that our models change when atoms come together to form molecules (although a molecular orbital theory is beyond this material, you can set the stage for its later acceptance).

CHAPTER NINE

1. In having the students answer questions such as this one, it is important to not allow them to merely read the definition from the text but to clearly define the two ideas in their own words. This strategy prompts the students to again think about what is meant by the phrase “lower in energy,” as they did in Chapter 6.
2. Although this concept is discussed in the text, it is a good idea to have three-dimensional models for the students to manipulate as they consider the question. Once the students see the sources of the two pi orbitals and one sigma orbital, a lot of the confusion is lifted.
3. One reason to ask the students this question is to make them more critical readers. That is, they should pay attention to the graphs and figures and not just skip over them. This question is explained in full in Section 9.3 but having the students explain in their own words is a good idea, especially in this case. Students are often more concerned with the product than with the process, and this question (and Section 9.3) emphasizes the process of evaluating a model and making appropriate changes.
4. This question also focuses on the phrase “lower in energy.” If one electron in the sigma 1s molecular orbital was excited to the antibonding orbital, the H₂ molecule would not be energetically more favorable than two separate hydrogen atoms. Although this question is very theoretical, it might be worthwhile to listen to how the students respond.
5. The notion of hybridized orbitals is quite confusing to students, and it is a good idea to have them work together on at least one of these problems.
6. We can determine molecular structures, sometimes directly by X-ray diffraction (Chapter 10). Our view is that if electrons are in orbitals in atoms, then they must be in orbitals in molecules. However, atomic orbitals do not point in the right direction for some molecules. We therefore had to invent the concept of hybrid orbitals. We decide which hybrid orbitals are used once we know the structure. Structures are the facts, and hybrid orbitals come from a model to rationalize these facts.
7. It is often a good idea to have the students compare and contrast different models we use. Especially focus on why one model is preferred to another depending on what we are trying to explain. For example, the VSEPR theory is quite simple, but does not explain such phenomena as paramagnetism.
8. This relationship is presented in the text. As mentioned above, it is helpful to have the students actively explain figures and tables from the text. They should realize that bond order comes directly from the model, while the other two quantities may be determined empirically.

CHAPTER TEN

1. Instead of allowing the students to merely say “because of surface tension” as a complete answer to this question, have the students explain what they mean by surface tension, and even have them draw a magnified view of the water and the soapy water. This approach is consistent with earlier chapters, which stressed pictorial representations of molecules.
2. When several hundred UIUC students given this question, about 40% correctly chose “d.” However, “a,” “b,” and “c” are all popular (each chosen by about 20% of the students). It is helpful to get an idea about what the students think about physical equilibrium because many times these ideas resurface when they discuss chemical equilibrium. As with other questions, it is more important to listen to how the students defend their choices and critique the other options than to simply tell them the answer.
3. Many students think of the term “boiling point” as meaning the boiling temperature. This question is meant to prompt students to think about how dynamic physical (and eventually, chemical) processes are and about vapor pressure. If possible, it would help for the students to see this process in a class demonstration.
4. Yes, there are some substances in which only dispersion forces are present such as naphthalene and polyethylene which are solids at room temperature. The fact that these substances are solids at room temperature tells us that their interparticle forces are stronger than those of substances that are liquids at room temperature, such as water, in which there are hydrogen bonds.
5. The nature of the force stays the same. As the temperature increases and the phase changes, the changes solid → liquid → gas occurs, and a greater fraction of the forces are overcome by the increased thermal (kinetic) energy of the particles.
6. Students have difficulty with explaining what the term “vapor pressure” means, even though it is explained in the text. Make sure that the students explain this idea using their own words. Also, many students do not think of solids as having a vapor pressure.
- 7, 8. Many students will think that vapor pressure decreases over time as the water evaporates (less water means less vapor pressure). Also, some will not realize that the vapor pressure of water at 100°C is 1.0 atm.
9. The students should be able to explain this idea from their reading in the text, but make sure they explain it in their own words. This discussion will help them to become more active and critical readers.

10. Intramolecular forces will be greater if the molecule melts and boils before it decomposes. This question gets to students' understandings of these terms, as opposed to just their knowledge of the text definition.
11. This question is similar to the question, "Why do bonds form between atoms?", in that it gets at the "why?" instead of just the "what?" Asking these questions of students (and allowing the students to ask them of us) helps the students become more critical thinkers.

CHAPTER ELEVEN

1. Again, one reason to ask the students a question like this one is to make them critical readers. While this concept is explained in the book, it is a good idea to listen to the students as they try to explain it in their own words. This edition places an even greater emphasis on representing magnified views of molecules, and it is instructive to have the students be active participants in this process. This question provides a good opportunity to get the students thinking about what the term "equilibrium" means, which will be important in later chapters.
2. This problem is similar to question 1, but a bit more complex. In this case there is a mixture (albeit a small one) of the two liquids, and the vapor pressures eventually become equal such that solutions end up in both beakers. A discussion of a sufficiently theoretical question such as this one can be a good opportunity to better understand how the students are making sense of such concepts as "vapor pressure" and "equilibrium."
3. The emphasis here is on having the students draw pictures and discuss osmotic pressure in the context of a problem that they have not seen before. The students may know what happens (the freshwater plant shrivels and the saltwater plant expands and possibly bursts), but they probably have not thought about it using the ideas of osmotic pressure.
4. The answer to this question is explained in Section 11.4, but as with some of the other questions it is important to listen to the students as they explain the problem in their own words (especially the reasons why). Many times, but especially in this case, the concepts are quite difficult to understand, and students may have learned to be rather passive readers (and unfortunately many are passive listeners even at the most brilliant of lectures). Therefore it is important to have the students take a more active role as they discuss their ideas with each other.
5. The responses given in the question represent actual student responses when trying to explain freezing-point depression and boiling-point elevations. Interestingly, after discussing this problem for quite some time (over half an hour at times), the students often come to the conclusion that this explanation is as good as any. Finding out what the students like and dislike about this

explanation (as opposed to whether they merely agree with it) is important. With an emphasis on vapor pressure, many students will claim that a change in vapor pressure can account for these properties (although the freezing point is still tricky, especially since many students believe that the salt freezes in the water), but when asked why vapor pressure matters, many come back to this explanation.

6. This question is meant to get the students thinking about whether salt freezes in the water, and to see how/if they incorporate the idea of vapor pressure of a solid (ice) in their explanation. The goal here is not to directly tell them an answer, but to try to better determine the ideas they use in their reasoning. This question provides another opportunity to have the students draw their ideas.
7. If the students have already learned about phase diagrams, they can see that water is rather unique in that an increase in pressure can actually cause a solid to liquefy. Thus, if salt is placed on the road and pressure is applied (from a car, for example), the ice could liquefy, salt could go into solution, and the salt-water solution would be less likely to refreeze. However, this response is a rather simplistic answer to a very complex problem and should not be taken as “the answer.” As with the spirit of most of the Active Learning Questions, the goal here is to prompt students to verbalize their ideas.
8. There are many variables in this question, and it is therefore hard to settle on one definite answer (for example, one person may screw the top of the bottle more tightly; if a student suggests this point, it would be best not to simply disregard that suggestion, but to ask “Why would that matter?”). All other things being equal, though, your friend's drink would be more carbonated because of the smaller volume above the drink. Again, have the students explain their answers.
9. Students often see the sudden use of molality in the colligative properties equations as arbitrary. Understanding why it is used is helpful for their understanding of science in general as a form of logic rather than disparate facts.
10. In this case, the vapor pressure of a solution is dependent on its concentration. The concentration will increase over time as the water evaporates. Thus, the vapor pressure will decrease over time.

CHAPTER TWELVE

1. Students have difficulty with the difference between kinetics and thermodynamics, and this question is a good way to get them to distinguish between the two. Many times students equate a fast reaction with a thermodynamically stable one (or with one that has a large K value, or with spontaneity). It is therefore a good idea to have the students differentiate among kinetics, thermodynamics, equilibrium, and spontaneity.

2. The rationale for asking this question is that many times students are not sure how the theory relates to the practice. One would hope that students, after reading the chapter and doing the problems could answer this question, but it doesn't generally seem to be the case. The text provides some examples, and if the students suggest these, have them explain the examples.
3. Students all too often merely memorize Table 12.6 without putting much thought into the conceptual differences among the half-lives. Answering this question allows them to explain reactions of different order and to think about what the term "half-life" means, and why it would differ for different orders of reactions.
4. As the temperature increases, the value for k should increase, although this increase is not linear. What is most important in this case is to have the students discuss why this is true, and it is up to you to decide how in-depth you wish to go with the Arrhenius equation.
5. As with other questions like this one, it is important to have the students decide what they like and do not like about the statements, especially what they do not like. These statements are actual student statements. While the rate does decrease because of the decreasing concentration of the reactants, the reaction need not "warm up."
6. The zero order reaction seems odd to many students. While most can do problems with it, many, when asked, admit they do not understand what it means for a reaction to be zero order. The text describes this concept, but the explanation is often missed by the students. As mentioned before, a potential byproduct of asking questions like this one is to inspire the students to become more active readers.
7. Recall from Chapter 3 that students may not be sure what a chemical equation represents, although they can balance the equations. Also, many times rate laws seem to have the same order as they would if taken from the balanced equations. Students are confused by this, especially by the difference between the steps in a mechanism and the balanced equation.
8. In this question, the student should explain why the half-life of zero-order reactions decreases with time, the half-life of first-order reactions is constant, and the half-life of second-order reactions increases over time. In doing so, they should achieve a better understanding of how concentration affects the rate of reaction for each order and what half-life means.

9. The rate constant is dependent on the nature of the reactants and the temperature. Concentration and order do not affect the constant. Students often confuse rate with the rate constant.

CHAPTER THIRTEEN

1. These questions were also asked of several hundred incoming UIUC students, and the results were quite interesting. In the case of adding chemical A to the flask, 85% of the students correctly claimed that the concentration of A would be higher once a new equilibrium was established, yet only 55% claimed that the concentration of B would be lower. In fact, 36% claimed that the concentration of B would remain unchanged, making one wonder how the reaction could proceed at all. At least half of the students claimed that the concentrations of C and D would change (most believed they would increase). Many students are confused, when directly asked, by the fact that adding only one reactant causes an increase in the amount of product; "But you are not adding any more B," they will claim. Also, many students seem to believe that the reactions are unidirectional, even though they will be able to solve the problems. Allow the students ample time to discuss their ideas.
2. Consistent with the increased emphasis on diagrams throughout the eighth edition, this question asks the students to graphically represent the product and is therefore a combination of qualitative and quantitative responses. Again, the goal is to get the students to better understand what is occurring with chemical equilibrium.
3. It may be worthwhile to have the students perform the calculation to check on their answer after discussing what they believe should happen (you would have to provide them with a value for the equilibrium constant, or you could ask them what information they would need to know to solve this problem). When the H_2 is added should not matter, but hearing what the students say is important.
4. In terms of increasing equilibrium concentrations of D, the answer is "ii," "i," and "iii." In terms of B, the answer is "iii," "i," and "ii." This question goes well with question 1 above.
5. For part a the range should be between 1 and 2 M (closer to 2 as K decreases); for b the range should be between 1 and 3 M (closer to 3 as K decreases); for c the range should be between 1 and 2 M (closer to 2 as K increases); and for d the range should be between 0 and 1 M (closer to 1 as K increases). Students have had quite a bit of difficulty with this one question in the past.
6. As noted in the discussion of question 1, this question has been asked by several students (and probably wondered but left unasked by many others). This question surfaces concepts that these students may be too shy to ask, and

allows them the opportunity to discuss it. Having the students draw pictures here would be helpful.

7. Students have believed this statement to be true, indicating that they believe that the most important factor is that K stays constant; that is, they do not necessarily consider the relevant chemistry (such as in question 5). In this case, the amount of C is impossible, but it allows for the “correct” K value. Again, it is important not to just treat this question as a “trick” and quickly tell the students the answer, but to allow them ample opportunity to discuss these issues themselves.
8. Le Châtelier’s principle is an example of a memorized rule that is very overused by the students. These questions allow students to see that an understanding of a concept cannot be achieved merely by memorizing a rule. The pressure is held constant in this problem by allowing the volume to change, which is also a perturbation of the system. This question is quite a complex problem, and you should expect the students to have difficulty with it. Also, if you choose to use this question, allow plenty of time for discussion because it could cause the students to become more confused about equilibrium. However, this potential drawback is outweighed by the potential benefits of having the students realize that it is not their job to memorize rules, but that science is about understanding.
9. Students are often confused between equilibrium condition and equilibrium constant. In this case, the answers are “c” and “d,” but many students believe that by changing the concentrations of the reactants or products, one changes the equilibrium constant.

CHAPTER FOURTEEN

1. Many students see the pH of water (or a neutral solution) defined as 7.00 regardless of the temperature and even though the pH scale is explained in the text. Although they read that the pH changes (decreases) as you heat water, they sometimes assume that this statement means that the water is no longer neutral, although it is.
2. Strength and concentration are confusing to students and often equated. HCl is always strong but its concentration can vary; along the same lines, ammonia is always weak.
3. The first graph seems to pose the more difficult challenge for the students because many claim that as an acid becomes more concentrated, it gives more H^+ , so the dissociation increases. This is true, but the question asks them to graph percent dissociation, which actually decreases. The text provides an explanation, but this is another concept that the students seem to need to discuss to more fully appreciate. Some students have even used applications of Le Châtelier’s Principle in explaining this problem.

4. The purpose of this question is to get the students to think qualitatively about major species and reactions. It is another good problem for which to have them draw pictures. It also provides an opportunity for students to differentiate what is occurring in solution with how they solve these types of problems (for example, many times we have them assume a reaction goes to completion given a sufficiently high value of K , so students sometimes mistakenly believe that these reactions can go to completion at our will).
5. This question is another example of a problem that has the students qualitatively explain something covered in the text.
6. This question is consistent with the increased emphasis on representing reactions in this eighth edition. With 10 molecules, you must be careful about the K for HCl and the relatively large value of K for the weak acid (many students will have 9 of the 10 molecules dissociate). For part D, from strongest to weakest base, the answers are A^- , H_2O , and Cl^- . Make sure that students explain their answers, even if they are correct.
7. This question is an actual UIUC student question. It has also come up when students are asked to calculate the hydroxide concentration of a solution with a pH less than 7. This question is intended to get at the idea that chemical equilibrium is a dynamic process.
8. Students often confuse the models that we use to think about reactions with the mathematics that we use to solve the problems. Many times we emphasize choice “a” mathematically, but tell them choice “c” as a model. It is important to see what they choose, why they choose it, and how they explain away the other choices.
9. This problem is like question 8; that is, it has the students differentiate between math and reality. In this case, the purpose is to have them realize that all reactions are happening at once, but that it is easier to calculate the results by making assumptions. It is up to you to decide whether you would like the students to actually solve the problem. What you may find is that even the students who are able to solve this problem may have difficulty answering these questions. Given the HCl concentration of 0.02 M, the pH would be 5.8.
10. The anion could not be strong, but many students may think that it could be if it comes from a weak acid; that is, often students have memorized the fact that the conjugate base of a weak acid is a strong base. The anion could be amphoteric if it comes from an acid such as phosphoric acid (for example, if the salt is NaH_2PO_4).

11. This question has confused students even after they have seen a mathematical derivation. Many also use this incorrect relationship. Again, it is important to bring up this idea because many students think it but remain quiet. Have the students actually go through the derivation and explain it.
12. This question is like question 6 in Chapter 3 in that it makes students think about what they need to solve a problem as opposed to trying to figure out how to “plug and chug” the numbers to arrive at a correct answer. One possible answer is that the students must know the dissociation constant values of HX and HY, and the acid with the lower dissociation constant would make the stronger (although still weak) conjugate base, and that salt would have the higher pH.
13. Make sure students explain what pH means in their own terms. The answer to the second part of this question is false (concentration plays a factor).
14. Another “why” question. Students can memorize this number, but they should also have an understanding of its derivation. It can lead to a discussion of the temperature dependence of pH.
15. The answer (yes) is easy enough to prove, but many students instinctively think the answer is no.
16. Many students memorize the statement “the conjugate base of a weak acid is a strong base.” This is presumably because students think of acids and bases as opposites and the terms weak and strong as opposites. The students should consider the reaction of a conjugate base of a weak acid with water, and think about what the K_b value would be before jumping to a conclusion. This is another example of students trying to memorize a simple statement as opposed to thinking about the situation.
17. Like question #5, this is to get the students to think about acids, bases, and salts. Even without doing calculations, the students should be able to order the chemicals according to acidity, and then match the pH's. As an extra, you can ask the students to determine the concentration of the chemicals, and to decide if they are all equal. In this problem (disregarding significant figures), the pH's are as follows: pH=1 for HBr, pH=2 for HF, pH=5 for HCN, pH=6 for NH_4F , pH = 6.5 for $\text{CH}_3\text{NH}_3\text{F}$, pH=8 for NaF, pH=11 for NaCN, pH=11 for NH_3 , pH = 13 for NaOH.
18. Another non-numerical (though arguably a semi-quantitative problem) that has the students think about the appropriate reactions. For BX to be acidic, B^+ should be a weak acid. While it is more likely for HX to be a strong acid (so that X^- is not basic), HX can be a weak acid, provided that B^+ is a stronger acid (although still weak) than X^- is as a base (although X^- would also be a weak base). Thus all of the choices could be true.

CHAPTER FIFTEEN

1. The fact that the concentration of baking soda does not change the pH should make for a good discussion. Have the students decide on the major reactions; this problem is also intended to get the students to think about major species. It may be a good idea to ask the students to draw pictures here as well. Note that HSO_4^- acts essentially exclusively like an acid, although HCO_3^- is amphoteric.
2. This question come from actual students and is a good one to ask if for no other reason than to unearth unasked questions. It also gets the students to consider that a buffered solution will change pH and reacts just as they could predict with major species, yet the ratio of conjugate base to acid stays relatively constant. This concept is one reason why the Henderson-Hasselbalch equation should be treated carefully, as students all too often just plug numbers into it with little consideration of the concepts involved. Many times students think of buffered problems as completely different from acid or base problems.
3. As with question 2, this question prompts students to think critically about buffered solutions to the extent of how they work and what makes one better than another. There is an extensive discussion of this idea in the text, but asking the students to explain a question such as this should shed some light on the students' ideas, and should help to make the students more critical and more active readers. If we mix twice as much acetic acid as NaOH, we will have acetic acid and the acetate ion in solution, which is a weak acid and its conjugate base.
4. We cannot make a buffered solution with HCl and NaOH since HCl is strong and so its conjugate base is very weak (for all practical purposes, it is neutral).
5. Another example of asking students to discuss a question discussed in the text.
6. A good follow-up question to number 5. Students will often memorize how to calculate pH values at various points but not understand why. Give the students the opportunity to consider why the formulas they use work.
7. The major species are HA, H^+ , Cl^- , and H_2O . Depending on the amount of H^+ in solution, we may not have to solve the equilibrium with HA and water. That is, the H^+ from the HCl may dominate. The pH of just HA will be higher.
8. The major species are HA, Na^+ , A^- , and H_2O . We need to know the amounts and the K_a value of the acid HA to solve this problem. Since A is a conjugate base of HA, the pH of the mixture will be higher than that of just HA.

CHAPTER SIXTEEN

1. Only “c” will affect the total amount. The others will affect the rate at which they dissolve, but not the amount.
2. Another example of relating theory and practice; it is good practice for the students to try to develop experimental verification on their own. This type of question also allows you to better assess their understanding of K_{sp} .
3. Students often become confused about what K_{sp} values actually measure and how they are related to solubilities. If a K_{sp} value were zero, it could mean that the solid does not dissolve into ions at all (the product of the ions is zero).
4. This question represents an addition to question 2 in that many students will equate K_{sp} values with solubilities such that a higher K_{sp} necessarily means a greater solubility. While the two are related, it is important for the students to understand that they are not the same thing and to be able to explain why they are different (the number of ions is important in this relationship).
5. This question is related to both questions 3 and 4. In this case, the students should be asked to explain the underlying theory of this before they start looking up numbers (many of them, if allowed to look up numbers immediately, would be more interested in trying to plug them into an equation somehow than in thinking conceptually about the problem). In its use of K_{sp} values, though, this is a good introductory problem in having the students calculate solubilities from K_{sp} values, and in thinking about what solubility means.
6. In general, the K_{sp} of a solid will increase with increasing temperature. However, this situation is not always the case and you must decide in how much detail you wish to discuss this issue with your students (the fact that some solids become less soluble with an increase in temperature of the solvent).
7. Because the sulfide ion comes from a weak acid, the sulfide ion has more affinity for the hydronium ions than does the chloride ion. This being the case, the silver sulfide is much more soluble in an acidic solution than in water. The solubility of silver chloride is not affected. Use this problem to stress the importance of major species and common ions (make the acidic solution be a solution of HCl and ask students to redo the problem), and as an introduction to multiple equilibria.
8. The students need to understand what is meant by K_{sp} and K_a to answer this question. AgY is more soluble in an acidic solution.

CHAPTER SEVENTEEN

1. In this case, in terms of positional randomness, the forward process (liquid to gas) is favored. In terms of energy randomness, the reverse process is favored (because going from gas to liquid would release energy into the surroundings).

This question should prompt the students to differentiate between the surroundings and the system.

2. The change in entropy of evaporation should be larger. As with other questions, have the students explain their answer and use their discussions to better understand their ideas (in this case, specifically about what is meant by ΔS).
3. As mentioned earlier, students have difficulty with the concept of “lower in energy.” In this case, the reaction is exothermic (ΔS_{surr} is positive), but ΔS cannot be determined with the given information.
4. Another example of having the students relate theory and practice.
5. This concept is explained in the text, but should be discussed by the students. This question was asked by a student at UIUC, and many other students agreed with him (although it is not a correct statement). It is another example of a problem that should elicit ideas that the students sometimes believe but do not ask.
6. The focus here is on not allowing the students to merely memorize the equations, but to critically think about why the equations are what they are. It is important (and not always obvious to the students) that students recognize the rationale for the equations.
7. For “a,” entropy increases (+) since a vapor is produced from a liquid. For “b,” entropy is decreased (-) since ice is more ordered than liquid water. For “c,” entropy decreases (-) since the gas is in a smaller volume. For “d,” entropy increases (+) since the particles are in a larger volume and “spread out.”
8. ΔS_{surr} is dependent on heat flow to the surroundings. Thus for exothermic reactions, ΔS_{surr} is favorable, and for endothermic reactions it is unfavorable.
9. This is another semi-quantitative problem requiring the student to think about a problem without merely plugging numbers into an equation. Because this process is carried out above 100°C , it is spontaneous (thus ΔS_{univ} is positive). Since a liquid is changed into a gas, ΔS is positive. Because the process is endothermic, ΔS_{surr} is negative.
10. If the reaction is exothermic, high temperatures will favor the reverse reaction.

CHAPTER EIGHTEEN

1. This question serves as a good introductory problem in that the students should conceptually explain the workings of a cell before doing any calculations. Again, this idea is explained in the text, but this fact does not guarantee that the students are actively reading the explanation. Having them explain figures in the text is a good way of assuring that the students read the book critically.
2. A good follow-up question to question 1, and a good chance for you to better understand the students' ideas. The text explains this idea, although not explicitly. This question requires the students to synthesize a lot of the information that they receive from the text and the lectures.
3. In this case, zinc should be used. This problem requires the students to use Table 18.1 (standard reduction potentials) although this fact is not mentioned in the question. Allow the students some time to figure out what they need to know to solve this problem.
4. Many times students have seen a demonstration in which a penny is placed in nitric acid (it may be a good idea to do this demonstration for the students—but provide ample ventilation!). Using the values in Table 18.1, the students should be able to figure this answer out, and this problem gives them the opportunity to figure out how to use these values (when to reverse the reactions, whether to multiply the potentials, and so forth).
5. In this case the voltage should be 0.29 V, and the electrons should flow from the Cr (anode) to the Fe (cathode). The equation is $3\text{Fe}^{2+} + 2\text{Cr} \rightarrow 3\text{Fe} + 2\text{Cr}^{3+}$. While there are other, similar end-of-chapter problems, it may be a good idea to have the students work together the first time they attempt such a problem.
6. Some students believe that if F_2 (for example) is a good oxidizing agent (it is the best one listed in Table 18.1), then F^- is a good reducing agent. Just from the data in the table, nothing can be said quantitatively about the reducing potential of the species F^- (although we could easily see that F^- would not want another electron). The important point here is for the students to realize what they can and cannot tell from the given information.
7. A good follow-up to question 6. We cannot tell anything about A^+ and B^+ as reducing agents, but we can say that B^+ is a better oxidizing agent.
8. This question gives you a good opportunity to better understand the students' ideas concerning ΔG , w , cell potential, and Q . It also serves as a good introduction to the concept of a concentration cell.
- 9, 12. These questions are related ($\Delta G = -nFE$). By solving for ΔG in question 12, the first two potentials can be used to calculate the third.

10. This question is a good chance for students to reflect after they have attempted a few calculation problems with galvanic and concentration cells. You may find that even the students who successfully solved the problems are not actively thinking about the issues involved. Stress to the students that by considering these types of questions, they are less likely to need to resort to merely memorize equations.
11. A good follow-up to question 10 (especially to see how the students believe E° is affected, which it should not be). The cell voltage should be increased by an increase in silver ion concentration and decreased by an increase in zinc ion concentration.
13. The answer to this is “No.” This question is to allow the students to see that the value of 0.00 V for the standard hydrogen electrode is arbitrarily set.
14. This is false. Along with this answer, have the students explain why a concentration cell “works” at all. The students often don't consider what drives such a cell.

CHAPTER TWENTY-ONE

1. This question is a good introductory problem to introduce to students the ideas of coordination compounds and to have them name these compounds. The formula should be $K_2[PtCl_6]$, and the name is potassium hexachloroplatinate(IV).
2. This question requires the students to use their knowledge to solve a novel problem and synthesize some of the information they have learned. Also, it is not specified precisely in the problem how to go about solving it, so the students have to figure the procedure out on their own. In both cases, the nickel ion has a charge of +2 thus there are 8 d orbital electrons. The paramagnetic complex is tetrahedral, and the diamagnetic complex is square planar.
3. Both iron ions have a charge of +2 (thus allowing for 6 d orbital electrons). The only difference in this case are the ligands. Because H_2O is more of a weak field ligand than CN^- , there is more of a chance that Δ is smaller, making $Fe(H_2O)_6^{2+}$ more likely to be paramagnetic. That is not to say that it is paramagnetic, but it is more likely to be paramagnetic. This question requires quite a bit of thinking on the part of the students, and ample time should be given to them to discuss it.
4. In comparing the high-spin case to the low-spin case, you can see that the conditions given in the problem will be satisfied if there are 4 d orbital electrons or 7 d orbital electrons. There can be many answers that satisfy these requirements.

PART VI: DISCUSSION OF THE *INQUIRY BASED LEARNING GUIDE*

PREFACE

Inquiry learning is based on the premise that students will better understand the material if they are active participants in the learning process, not just passive sponges for information. Dictionary definitions of inquiry include aspects such as questioning, a search for knowledge, deliberation, and interrogation. Basically, to inquire is to ask and consider questions, and this requires the students to be engaged in the material.

The questions and activities in the *Inquiry Based Learning Guide (IBLG)* are designed to get students to consider the underlying concepts involved in understanding chemistry. We want the students to recognize that learning chemistry goes beyond “getting the right answer” to algorithm-based exercises. We want them to learn to think like chemists—to be able to solve problems because they truly understand the underlying concepts, not because they have memorized a solution to a particular type of problem.

Your teaching experience has probably shown (and there has been a great deal of research to back this up) that students can be successful at solving even relatively difficult math-based chemistry problems yet still not have a firm understanding of the chemical principles involved in the problem. This is partly due to the fact that students generally solve these problems by using memorized rules, algorithms, or equations. While we as instructors mean for the problems to be conceptual in nature, students often exchange thinking about the chemistry with finding a short cut to solving a specific problem. Because of this, many of the questions in the *IBLG* explicitly ask the students to explain a term, equation, rule, or algorithm. In other words, we ask the students to explain what we have long been assuming they had to know to solve the traditional problems.

HOW DOES THE *IBLG* HELP STUDENTS LEARN CHEMISTRY?

There are two main impediments to learning that the *IBLG* addresses. The first consists of the content misconceptions, and the second is the way in which many students define teaching and learning.

The questions and activities in the *IBLG* have been specifically written to surface students' initial ideas and content misconceptions. Students (and instructors) often do not realize the misconceptions harbored by the students until the students have to explain their ideas about specific topics. Once these misconceptions come to light, they can be confronted, and conceptual change can take place.

While these content misconceptions undoubtedly play a role in students' lack of understanding, there is an even more fundamental difficulty: misconceptions about what it means to really understand. For example, we have found:

1. **Students are passive readers of textbooks.** They often view the textbook as simply a place to find their homework assignments, and perhaps to look for sample problems and solutions. Even students who take the time to read some of the text will often skip over graphs and figures, which are crucial aids for increasing conceptual understanding. Many of the questions in the *IBLG* send the students back to the text to consider specific statements, or to discuss the significance of a table, graph, or figure.
2. **Even successful students rely more on math ability than problem solving approaches.** Although at the University of Illinois we have some of the brightest math students in the country, many of these students are not very proficient at thinking through problems in chemistry. We have had many students making claims such as “Once I am shown how to set up a problem I can usually solve it,” and genuinely believe that this demonstrates understanding. It is not so much that they are trying to get out of thinking, but that they are defining understanding in terms of algorithms. If we give them a problem without an obvious algorithm, they struggle. The problems in the *IBLG* ask the students to consider the chemistry in the system, and to reflect on how they solve the problems.

HOW DO THE QUESTIONS IN THE *IBLG* ADDRESS THESE CONCERNS?

As stated previously, the essence of the questions in the *IBLG* is to ask the students explicitly to discuss the fundamental ideas of chemistry and to consider how they solve problems. These problems generally require the students to explain rather than calculate; even when the students are asked to calculate, they are also asked to explain their reasoning and the significance of the answers. In this way, the students are forced to confront misconceptions, and they learn to actively seek answers (the nature of inquiry) and evaluate their problem solving abilities. Let's consider a few sample problems from the *IBLG* and discuss their rationale.

Example #1: Relating microscopic to macroscopic

Question: Provide microscopic drawings down to the atoms for Figure 1.16 in your text.

Rationale: This problem not only sends students to the text, it also has the students consider their understandings of concepts such as *element*, *compound*, and *mixture* and relate them to the microscopic world. One of the reasons chemistry is so difficult for students is that we live in the macroscopic world, but discuss chemistry in the microscopic world. It is a good idea to get the students to relate these two as early and often as possible.

Example #2: Explicitly considering a term

Question: Why is the term “sodium chloride molecule” incorrect but the term “carbon dioxide molecule” is correct?

Rationale: This question has the students consider what the term *molecule* means and they must decide when it does apply and when it does not. It also addresses a misconception because since we can write the formula NaCl, many students believe it is a molecule.

Example #3: Explicitly considering a procedure

Question: On page 92 of the text, in describing how to determine a formula, it states the “easiest way to do this is to work with 100.00 grams of the compound.”

- Why is this the “easiest way”?
- Could we assume any mass? Prove your answer using a homework problem.
- Could we assume 1.0 mole of the compound? Support your answer.

Rationale: This problem not only sends students to the text, it also has the students explain why it is they solve a particular problem in a certain way, and consider if this is the only way to solve it.

Example #4: A math problem with no numbers

Question: Estimate the mass of air at normal conditions that takes up the volume of your head. Provide support for your answer.

Rationale: This problem requires the students to determine (and estimate) the data required to solve a problem, and then to solve it. This requires the students to explicitly think about the approach they will use to solve the problem. The goal is to help the students develop their own strategies for solving the problems.

Example #5: What if...?

Question: What if energy was not conserved? How would this affect our lives?

Rationale: This problem requires the students to discuss the significance of an idea (in this case energy conservation) by imagining what would happen if it were not true. In framing a question this way, the students are required to think more deeply about the topic.

HOW SHOULD THE *IBLG* BE USED?

A crucial part of learning involves true discussion of the concepts using questions that require the students to probe their understanding of the concepts. As instructors we know that trying to teach something to someone is one of the surest ways to expose our misunderstandings of the underlying concepts. Peer teaching is one of the best features of group discussions. However, for group work to yield real value, the questions must be carefully chosen. For example, while relatively traditional quantitative problems have their place in a chemistry course, they usually do not elicit much in the way of discussion. The main goal of traditional problems is usually to get the correct answer (although the process is also important, the emphasis is usually on the answer). That is not to say that one type of problem is more important than the other, but that both types need to be provided to the students. The questions in the *IBLG* come from our own experiences with increasing the level of discussion and peer teaching at the University of Illinois.

The ideal way to use these questions is to have relatively small groups of students (3 or 4) working together during class time. This serves many purposes including:

1. **The students are effectively modeling how to think about problems—** many ideas surface, debates occur, students often go down “wrong paths,” and they sometimes discuss important issues which are not explicitly written into the questions. This type of environment also teaches the students the benefits of actively thinking through a problem.
2. **Students are required to be active participants in their own education.** This is important for the reasons already mentioned, and additionally—we have found that many times students need to say something out loud before they can really evaluate it. For example, a student who believes he/she understands a concept (but does not) will often start out speaking quickly and confidently and then begin to trail off as he/she “listens” to what he/she is saying.
3. **You can interact with the students and question them further.** This is important to keep students from using short cuts to just get a correct answer, and to question the students even when they are correct (to keep them from being passive acceptors of answers). The answers we accept from the student are as important as the questions we ask of them.
4. **You get to know the students better more quickly.** This includes a better understanding of how they approach problems, their misconceptions, and group dynamics.

A FINAL NOTE OF CAUTION AND ENCOURAGEMENT

We hope you will find the questions and activities in the *Interactive Based Learning Guide* useful in your courses. That said, while having the students take a more active role in the learning process is a long term goal, the students often struggle with this initially. Students, especially those who have been successful algorithm solvers in the past, will often become easily frustrated with having to explain their ideas and their approaches to problem solving. However, the more you can show the growth in learning that these students experience, the more they will accept that this approach is a better one. One of the greatest compliments we have heard from one of our students is “I am trying to learn how to ask myself the same questions that you ask us. Then I will not need you anymore.” This is the true essence of life-long learning and surely a goal of any teacher.