CHEMISTRY

A MOLECULAR APPROACH

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Westmont College

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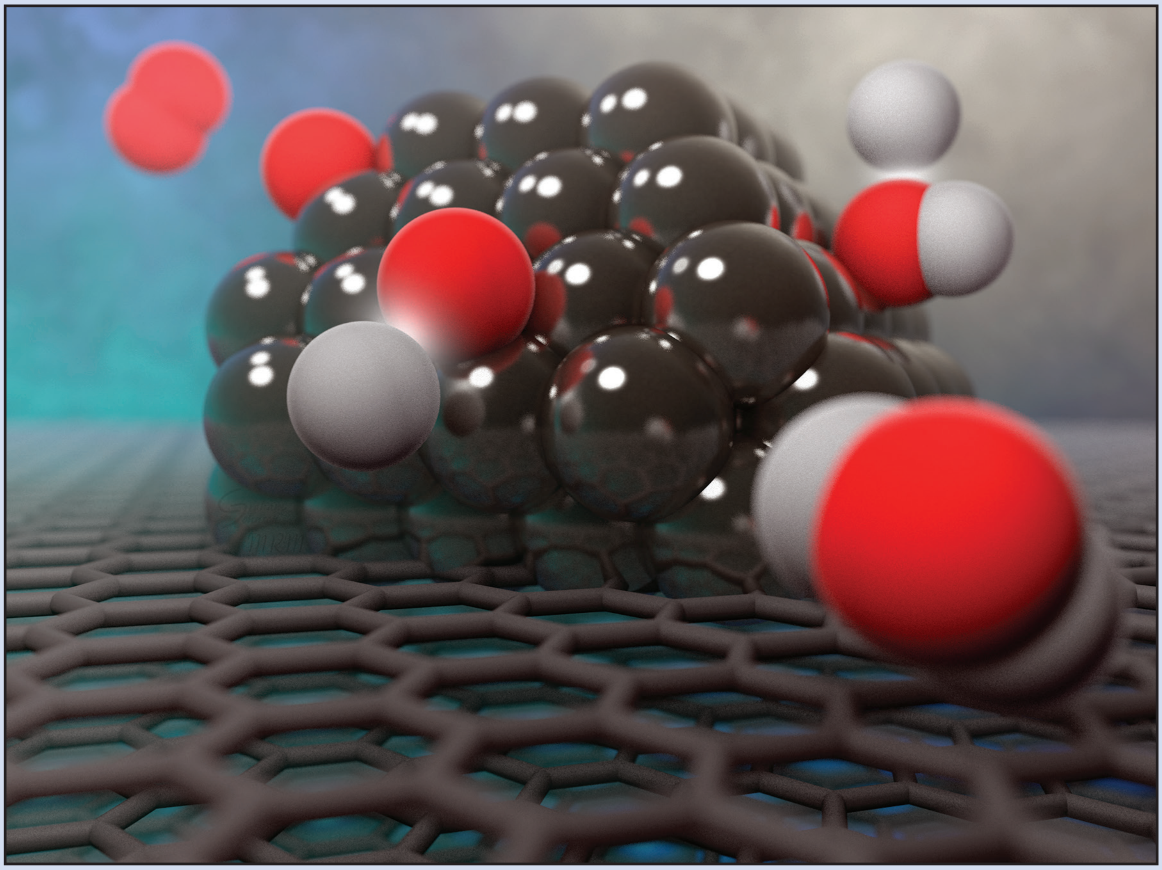

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PEARSON

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To Michael, Ali, Kyle, and Kaden
—Nivaldo Tro

To Cailyn, Carter, Colton, and Chloe
—Travis Fridgen

To Calvin, Nathan, Alexis, and Andrew
—Lawton Shaw

About the Authors

Nivaldo Tro is a Professor of Chemistry at Westmont College in Santa Barbara, California, where he has been a faculty member since 1990. He received his Ph.D. in chemistry from Stanford University for work on developing and using optical techniques to study the adsorption and desorption of molecules to and from surfaces in ultrahigh vacuum. He then went on to the University of California at Berkeley, where he did postdoctoral research on ultrafast reaction dynamics in solution. Since coming to Westmont, Professor Tro has been awarded grants from the American Chemical Society Petroleum Research Fund, from Research Corporation, and from the National Science Foundation to study the dynamics of various processes occurring in thin adlayer films adsorbed on dielectric surfaces. He has been honoured as Westmont's outstanding teacher of the year three times and has also received the college's outstanding researcher of the year award. Professor Tro lives in Santa Barbara with his wife, Ann, and their four children, Michael, Ali, Kyle, and Kaden. In his leisure time, Professor Tro enjoys mountain biking, surfing, reading to his children, and being outdoors with his family.

Travis Fridgen is currently Professor and Head in the Department of Chemistry at Memorial University of Newfoundland in St. John's, Newfoundland and Labrador. His research group studies the energetics, reactions, and structures of gaseous ion self-assembled complexes composed of metal ions and biologically relevant molecules such as DNA bases, amino acids, and peptides using a combination of mass spectrometry, tunable infrared lasers, and computational chemistry. Their research is aimed at answering fundamental questions such as why K^+ is associated with guanine quadruplexes such as telomeric DNA. He graduated with a B.Sc. (Hons) in chemistry from Trent University and a B.Ed. from Queen's University. His Ph.D. in physical chemistry is from Trent and Queen's Universities, where he studied the spectroscopy of reactive species in a cryogenic matrix environment. During his postdoctoral fellowship at the University of Waterloo, he first began conducting research using mass spectrometric methods. During a brief period as an assistant professor at Wilfrid Laurier University, he initiated a collaboration with a group of researchers from France to spectroscopically determine structures of gas phase proton-bound dimer ions. He teaches courses in physical chemistry, but he has mostly taught first-year chemistry courses (at Trent, Waterloo, Laurier, and Memorial). He lives in Mount Pearl, Newfoundland and Labrador, with his wife, Lisa, and four children, Cailyn, Carter, Colton, and Chloe. They are all avid fans of the Ottawa Senators and enjoy busy, active lives that include outdoor activities such as shovelling snow (good old Newfoundland).

Lawton Shaw received his Ph.D. in chemistry from the University of Calgary, in the area of photochemical reaction mechanisms of organometallic complexes. Shortly after graduating, he joined the full-time teaching faculty at Mount Royal College in Calgary, where he developed one of the first science courses at Mount Royal delivered partially online. This work led to a serious interest in online and distance education. In 2005, he joined the Centre for Science at Athabasca University, where he teaches and coordinates distance-delivered chemistry courses. This experience led to the book *Accessible Elements: Teaching Science Online and at a Distance,* which he co-edited. His research interests are split between the realms of teaching/education and environmental chemistry. He studies the effects of pharmaceuticals and personal care products on biofilms in freshwater ecosystems. He is a former president of College Chemistry Canada. He lives in St. Albert, Alberta, with his wife, Tanya, and their four children. Their family leisure time is filled with activities such as cross-country skiing, swimming, and camping.

Brief Contents

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Preface

To the Student

As you begin this course, think about your reasons for enrolling in it. Why are you taking general chemistry? Why are you pursuing a university or college education at all? If you are like most students taking general chemistry, part of your answer is probably that this course is required for your major or you are pursuing your education so that you can get a job some day. Although these are both good reasons, we think there is a better one. The primary reason for an education is to prepare you to *live a good life*. You should understand chemistry—not for what it can *get* you—but for what it can *do* for you. Understanding chemistry is an important source of happiness and fulfillment.

Understanding chemistry helps you to live life to its fullest for two basic reasons. The first is *intrinsic*: through an understanding of chemistry, you gain a powerful appreciation for just how rich and extraordinary the world really is. For example, one of the most important ideas in science is that **the behaviour of matter is determined by the properties of molecules and atoms**. With this knowledge, we have been able to study the substances that compose the world around us and explain their behaviour by reference to particles so small that they can hardly be imagined. If you have never realized the remarkable sensitivity of the world we *can* see to the world we *cannot,* you have missed out on a fundamental truth about our universe. The second reason is *extrinsic*: understanding chemistry makes you a more informed citizen—it allows you to engage with many of the issues of our day. Scientific literacy helps you understand and discuss in a meaningful way important issues from the development of the oil sands in Alberta (Chapter 6) to how the production of pharmaceuticals and personal care products affects our environment and our bodies (Chapter 12). In other words, understanding chemistry makes *you* a deeper and richer person and makes your country and the world a better place to live. These reasons have been the foundation of education from the very beginnings of civilization.

So this is why we think you should take this course and why we wish you the best as you embark on the journey to understand the world around you at the molecular level. The rewards are well worth the effort.

The Strengths of *Chemistry: A Molecular Approach*

Chemistry: A Molecular Approach is first and foremost a *student-oriented book*. The main goal of the book is to motivate students and get them to achieve at the highest possible level. As we all know, many students take general chemistry because it is a requirement; they do not see the connection between chemistry and their lives or their intended careers. *Chemistry: A Molecular Approach* strives to make those connections consistently and effectively. Unlike other books, which often teach chemistry as something that happens only in the laboratory or in industry, this book teaches chemistry in the context of relevance. It shows students *why* chemistry is important to them, to their future careers, and to their world.

Second, *Chemistry: A Molecular Approach* is a *pedagogically driven book*. In seeking to develop problem-solving skills, a consistent approach is applied (Sort, Strategize, Solve, and Check), usually in a two- or three-column format. In the two-column format, the left column shows the student how to analyze the problem and devise a solution strategy. It also lists the steps of the solution and explains the rationale for each one, while the right column shows the implementation of each step. In the three-column format, the left column outlines the general procedure for solving an important category of problems that is then applied to two side-by-side examples. This strategy allows students to see both the general pattern and the slightly different ways in which the procedure may be applied in differing contexts. The aim is to help students understand both the *concept of the problem* (through the formulation of an explicit conceptual plan for each problem) and the *solution to the problem.*

Third, *Chemistry: A Molecular Approach* is a *visual book.* Wherever possible, images are used to deepen the student's insight into chemistry. In developing chemical principles, multipart images help to show the connection between everyday processes visible to the unaided eye and what atoms and molecules are actually doing. Many of these images have three parts: macroscopic, molecular, and symbolic. This combination helps students to see the relationships between the formulas they write down on paper (symbolic), the world they see around them (macroscopic), and the atoms and molecules that compose that world (molecular). In addition, most figures are designed to teach rather than just to illustrate. They include annotations and labels intended to help the student grasp the most important processes and the principles that underlie them. The resulting images are rich with information but also uncommonly clear and quickly understood.

Fourth, *Chemistry: A Molecular Approach* is a *"big picture" book*. At the beginning of each chapter, a short paragraph helps students to see the key relationships between the different topics they are learning. A focused and concise narrative helps make the basic ideas of every chapter clear to the student. Interim summaries are provided at selected spots in the narrative, making it easier to grasp (and review) the main points of important discussions. And to make sure that students never lose sight of the forest for the trees, each chapter includes several *Conceptual Connections,* which ask them to think about concepts and solve problems without doing any math. The idea is for students to learn the concepts, not just plug numbers into equations to churn out the right answer.

Finally, *Chemistry: A Molecular Approach* is a book that delivers the depth of coverage faculty want and students need. We do not have to cut corners and water down the material in order to get our students interested. We simply have to meet them where they are, challenge them to the highest level of achievement, and then support them with enough pedagogy to allow them to succeed.

The Canadian Edition

Chemistry: A Molecular Approach, by Nivaldo J. Tro, is widely used in general chemistry courses at colleges and universities across North America. So, why do we need a Canadian edition? The short answer is that general chemistry courses in Canada are different from those in the United States. First-year chemistry curricula in Canada are generally at a higher level than what is seen south of the border. There is a need for a strong chemistry textbook that serves Canadian general chemistry courses.

The Canadian adaptation of *Chemistry: A Molecular Approach* drew very heavily on feedback from professors and instructors across Canada. As the Canadian authors, we took the reviews and consultations very seriously and did our best to adapt Tro's textbook accordingly. In general terms, the adaptation involved making the following changes.

International Conventions on Units, Symbols, and Nomenclature The field of chemistry is communicated according to conventions that are determined by the broader international chemistry community, through the International Union of Pure and Applied Chemistry (IUPAC). IUPAC continually releases recommendations on chemical nomenclature, definitions, symbols, and units. IUPAC recommendations are not static; they may evolve over time as new information comes to light. Although many textbooks state that they follow the recommendations of the IUPAC, you will find that the Canadian edition of *Chemistry: A Molecular Approach* scrupulously follows IUPAC recommendations for chemical names and symbols, nomenclature, and conventions for symbols and units in measurements. In the case of chemical nomenclature, there are a number of non-IUPAC chemical names that are so common that we have to include them along with the IUPAC recommended name.

S.I. units of measurement are used exclusively. Imperial units such as the gallon, pound, and the Fahrenheit scale of temperature have not been used in modern science for over a generation. IUPAC recommended defining standard pressure as 1 bar (or 100 kPa) back in 1982. This is the standard that has been adopted by chemists worldwide and is almost exclusive in second-year physical chemistry texts. Only in first-year textbooks does the atmosphere still linger as standard pressure. In this text, standard pressure is the IUPAC-recommended bar. Students will see pressure in various units, but we make little use of the atmosphere. When dealing with ideal gases, the most common value of R is 0.08314 L bar mol^{-1} K^{-1}.

In thermodynamics, we have adopted the recommended notation for enthalpy, entropy, and Gibbs energy changes, placing subscripts for changes after the delta sign rather than after H, S, or G. For example, the standard reaction enthalpy is expressed as $\Delta_r H^\circ$ rather than ΔH°_{rxn}. This is a subtle change that matters. The type of change (Δ) is marked on the Δ symbol (reaction, Δ_r; formation, Δ_f; and so on), rather than the type of thermodynamic quantity. We understand that this notation is not used everywhere. However, we believe that students should use standard notation throughout their education. Students who continue in chemistry or other sciences will eventually come across the standard notation in physical chemistry textbooks and in places like the *CRC Handbook of Chemistry and Physics* and the NIST Chemistry Webbook (http://webbook.nist.gov/). Furthermore, thermodynamic quantities like $\Delta_r H^\circ$ are always molar quantities and have the units kJ mol^{-1}, as recommended by IUPAC. Exclusive use of IUPAC-recommended units keeps students from getting into unit troubles when doing thermodynamic calculations.

Explicitly, we have provided the distinctions and connections between the unitless thermodynamic equilibrium constant, K_{eq} or simply K, and the phenomenological equilibrium constants, K_c and K_P, which can have units in terms of concentration and pressure, respectively, again in accordance with IUPAC recommendations. This is done in the most basic of terms, assuming that gases and solutions are ideal so that their partial pressures and concentrations are assumed to be numerically equivalent to their activities, setting up for a more rigorous treatment in second-year analytical and physical chemistry courses.

Following recommendations set out by the IUPAC ensures that we speak a common language—and teach a common language. Otherwise, students who go on in chemistry have to convert from the language learned in first year as soon as the very next year, when they take their first physical chemistry course.

Current Theories We have updated the text so that the most current, consensus scientific view is described. This is most notable in the case of bonding theory and the so-called expanded octet. In this case, evidence shows that the *d* orbitals have a negligible contribution to bonding, which means that full sp^3d and sp^3d^2 hybridizations should no longer be included in bonding theories, even though this idea continues to appear in general chemistry textbooks. This Canadian edition reflects the most current understanding of chemical phenomenon, at the first-year level.

Organic Chemistry The coverage of organic chemistry has been expanded to two chapters, reflecting the curricula in many Canadian universities, which provide additional organic chemistry coverage in first-year chemistry. The first organic chemistry chapter covers structure and bonding, stereochemistry, and structure determination. The second chapter covers organic reactivity, and it is organized according to reaction mechanisms.

Canadian Context Naturally, a Canadian edition will include Canadian examples. In some places, the Canadian content is fun, like the hockey goalie's "Quantum mechanical five hole" in Chapter 7. In other places, Canadian chemistry examples are serious and important, like the chemistry of the oil sands. Wherever Canadian content appears in this edition, it is there to promote student engagement. This book is meant for the Canadian student.

End-of-Chapter Problems One of the first things that professors consider when choosing a chemistry textbook is the quality of end-of-chapter problems. This is because, to learn chemistry, students need to work through meaningful exercises and problems. Tro's *Chemistry: A Molecular Approach* has extensive, high-quality problems.

First-year chemistry courses are perhaps the most important courses in chemistry programs, because they lay the foundation for all higher level courses. First-year courses introduce students to the language and discipline of chemistry, and some concepts are not touched on again in the entire undergraduate curriculum. Indeed, many Ph.D. comprehensive questions fall back to ideas learned in first year. This book was prepared with the full undergraduate curriculum in mind. If you are a student, we hope that the Canadian edition of *Chemistry: A Molecular Approach* helps you succeed in chemistry. We encourage you to make use of all of the features in this book that are designed to help you learn. If you are a professor, it is our hope that this textbook provides you with the strong content you need to teach first-year chemistry in a way that is true to our discipline.

Second Canadian Edition

For the second Canadian edition, our goal was to fine-tune the content from the first Canadian edition and ensure the concepts presented are in alignment with current, accepted theories. We also replaced or updated a number of the "Chemistry in Your Day" boxes to bring them up to date and add context for students. We have also added numerous electrostatic potential maps throughout the text to give students a visual aid to better understand chemical concepts related to electrostatic forces. Some of the substantial changes are described below.

Some material has been moved. For example, balancing redox equations has been moved from Chapter 18 (Electrochemistry) to Chapter 4 (Chemical Reactions and Stoichiometry); Chapter 9 (Chemical Bonding I: Lewis Theory) has been rearranged to make the content flow better for students.

In Chapter 7 (The Quantum Mechanical Model of the Atom), we revised and expanded the section on electron configurations to bring it in line with current literature, especially in regard to the transition metals. These changes should clear up some of the misconceptions that arise when students learn the *Aufbau* principle.

Chapter 10 (Chemical Bonding II: Molecular Shapes, Valance Bond Theory, and Molecular Orbital Theory) includes a section on larger conjugated systems as well as conductors, semiconductors, and insulators, which we feel is important for students who continue in chemistry. The molecular orbital diagrams in this chapter have also been updated, providing more "realistic" visual representations.

In Chapter 17 (Gibbs Energy and Thermodynamics), we added a more rigorous and chemically-relevant description of entropy and microstates which will benefit every student who continues in chemistry and other sciences.

Supplements

For the Instructor

MasteringChemistry® is the best adaptive-learning online homework and tutorial system. Instructors can create online assignments for their students by choosing from a wide range of items, including end-of-chapter problems and research-enhanced tutorials. Assignments are automatically graded with up-to-date diagnostic information, helping instructors pinpoint where students struggle either individually or as a class as a whole.

Instructor resources are password protected and available for download from the Pearson online catalogue at http://catalogue.pearsoned.ca/.

Instructor's Solutions Manual This manual contains step-by-step solutions to all complete, end-of-chapter exercises. The Instructor's Solutions Manual to accompany the second Canadian edition has been extensively revised and checked for accuracy. The Instructor's Solutions Manual can be downloaded from the online catalogue.

Instructor's Resource Manual Organized by chapter, this useful guide includes objectives, lecture outlines, and references to figures and solved problems, as well as teaching tips. The Instructor's Resource Manual can be downloaded from the online catalogue.

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PowerPoint® Presentations PowerPoint® lecture slides provide an outline to use in a lecture setting, presenting definitions, key concepts, and figures from the textbook. The textbook's worked examples and a selection of practice problems are also provided in PowerPoint® format. These PowerPoint® slides can be downloaded from the online catalogue.

Questions for Classroom Response Systems Another set of PowerPoint® slides provide sample exercises and questions to be used with Classroom Response Systems. These questions can be downloaded from the online catalogue.

Image Libraries All images, figures, and tables in the textbook are provided in PowerPoint® format. The images, figures, and tables are also available in a separate image library in jpeg or gif format. The Image Libraries are available through the online catalogue.

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For the Student

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Selected Solutions Manual This manual for students contains complete, step-by-step solutions to selected odd-numbered end-of-chapter problems. The Selected Solutions Manual to accompany the second Canadian edition has been extensively revised, with all problems checked for accuracy.

During the development of this book, we obtained many helpful suggestions and comments from colleagues from across the country. We sincerely thank the following instructors who were members of our Chemistry Advisory Board for this edition:

Phil Dutton, *University of Windsor*
Noel George, *Ryerson University*
Krystyna Koczanski, *University of Manitoba*
Andrew McWilliams, *Ryerson University*
Andrew Vreugdenhil, *Trent University*

We acknowledge Prof. Dietmar Kennepohl (Athabasca University) and Dr. Nicole Sandblom (University of Calgary), Dr. Neil Anderson (Onyx Pharmaceuticals), Drs. Chris Flinn, Bob Helleur, Karen Hattenhauer, Peter Warburton, and Chris Kozak, (Memorial University), Mr. Nicholas Ryan (Memorial University), and Drs. Lucio Gelmini and Robert Hilts (MacEwan University) for helpful discussions and insightful comments.

Dr. Ian Hunt of the University of Calgary worked with us in the early development of the organic chemistry chapters. He provided sage advice on the organization of these chapters and made numerous suggestions on how to present organic chemistry in a way that is both rigorous and accessible to the first-year student.

Professor François Caron of Laurentian University provided expert advice on revisions to Chapter 19, improving the presentation of nuclear reaction energetics so that it is consistent with the field of nuclear chemistry.

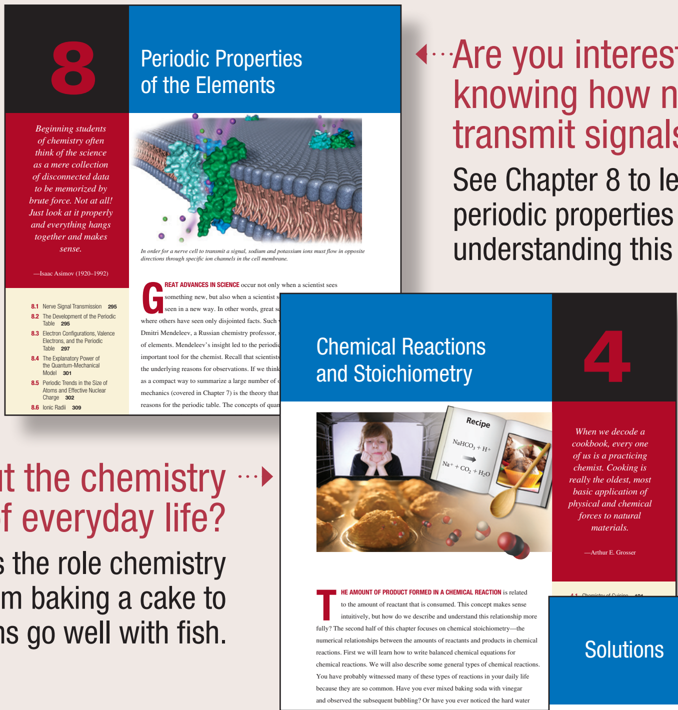
We would like to thank our wives Lisa and Tanya for their encouragement and their continuing patience during all the evenings and weekends we spent working on this book when we could have been with our families.

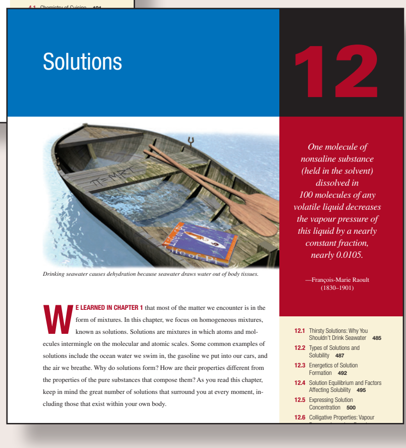
Finally, we would also like to acknowledge the assistance of the many members of the team at Pearson Canada who were involved throughout the writing and production process: Cathleen Sullivan, Executive Acquisitions Editor; Kim Teska, Senior Marketing Manager; Darryl Kamo, Program Manager; Martina van de Velde, Developmental Editor; Jessica Hellen, Project Manager; Anthony Leung, Senior Designer.

Travis D. Fridgen
Lawton E. Shaw

Relevant examples and clear language

Chemistry is relevant to every process occurring around you, at every second. The authors help you understand this connection by weaving specific, vivid examples throughout the text that tell the story of chemistry. Every chapter begins with a brief story that illustrates how chemistry is relevant to all people, at every moment.



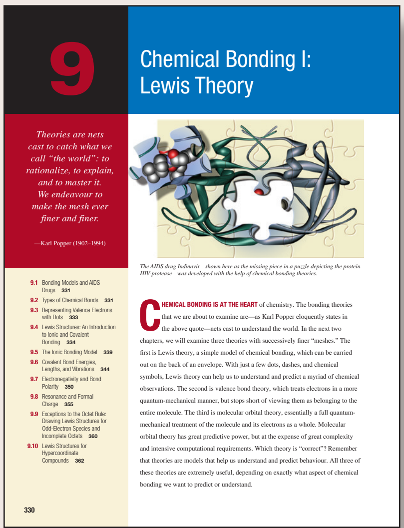


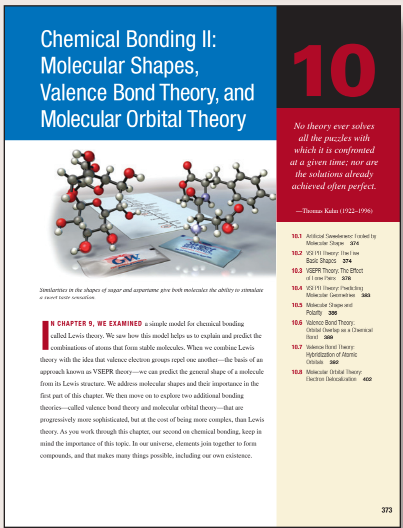
Are you interested in knowing how nerve cells transmit signals?

See Chapter 8 to learn why periodic properties are essential to understanding this process.

What about the chemistry of everyday life?

Chapter 4 illustrates the role chemistry plays in cuisine, from baking a cake to why lemons go well with fish.





These examples make the material more accessible by contextualizing the chemistry and grounding it in the world you live in.

connect chemistry to YOUR WORLD

Student Interest

Throughout the narrative and in special boxed features, interesting descriptions of chemistry in the modern world demonstrate its importance.

CHEMISTRY AND MEDICINE Bone Density

Osteoporosis—which means *porous bone*—is a condition in which bone density becomes too low. The healthy bones of a young adult have a density of about 1.0 g cm^{-3}. Patients suffering from osteoporosis, however, can have bone densities as low as 0.22 g cm^{-3}. These low densities mean the bones have deteriorated and weakened, resulting in increased susceptibility to fractures, especially hip fractures. Patients suffering from osteoporosis can also experience height loss and disfiguration such as dowager's hump, a condition in which the patient becomes hunched over due to compression of the vertebrae. Osteoporosis is most common in postmenopausal women, but it can also occur in people (including men) who have certain diseases, such as insulin-dependent diabetes, or who take certain medications, such as prednisone. Osteoporosis is usually diagnosed and monitored with hip X-rays. Low-density bones absorb fewer of the X-rays than do high-density bones, producing characteristic differences in the X-ray image. Treatments for osteoporosis include additional calcium and vitamin D, drugs that prevent bone weakening, exercise and strength training, and, in extreme cases, hip-replacement surgery.

Question

Suppose you find a large animal bone in the woods, too large to fit in a beaker or flask. How might you approximate its density?

▲ Magnified views of the bone matrix in a normal femur (left) and one weakened by osteoporosis (right).

▲ Severe osteoporosis can necessitate surgery to implant an artificial hip joint, seen in this X-ray image.

CHEMISTRY IN THE ENVIRONMENT Antifreeze in Frogs

Wood frogs (*Rana sylvatica)* look like most other frogs. They are a few inches long and have characteristic greenish-brown skin. However, wood frogs survive cold winters in a remarkable way—they partially freeze. In its partially frozen state, the frog has no heartbeat, no blood circulation, no breathing, and no brain activity. Within 1–2 hours of thawing, however, these vital functions return and the frog hops off to find food. How does the wood frog do this?

Most cold-blooded animals cannot survive freezing temperatures because the water within their cells freezes. As we learned in Section 11.9, when water freezes, it expands, irreversibly damaging cells. When the wood frog hibernates for the winter, however, it produces large amounts of glucose that is secreted into its bloodstream and fills the interior of its cells. When the temperature drops below freezing, extracellular body fluids, such as those in the abdominal cavity, freeze solid. Fluids within cells, however, remain liquid because the high glucose concentration lowers their freezing point. In other words, the concentrated glucose solution within the frog's cells acts as antifreeze, preventing the water within the cells from freezing and allowing the frog to survive.

▲ The wood frog survives winter by partially freezing. It protects its cells by flooding them with glucose, which acts as an antifreeze.

Question

The wood frog can survive at body temperatures as low as −8.0 °C. Calculate the molality of a glucose solution ($C_6H_{12}O_6$) required to lower the freezing point of water to −8.0 °C.

▲ Chemistry and Medicine boxes show applications relevant to biomedical and health-related topics.

◀ Chemistry in the Environment boxes relate chapter topics to current environmental and societal issues.

▼ Chemistry In Your Day boxes demonstrate the importance of chemistry in everyday situations.

CHEMISTRY IN YOUR DAY Oxygen in Spacesuits

As is shown in Table 5.3, the air we breathe contains 78% N_2 and 21% O_2. The other 1% is a mixture of Ar, CO_2, Ne, and other trace gases. Besides oxygen, human beings can survive without any of the other gases as they fill no physiological requirements. In the low pressures of outer space, pure oxygen is exactly what astronauts breathe. For the best mobility, it is an advantage to have as low a pressure as possible in a spacesuit when astronauts go on a spacewalk. In air at a pressure of 1 bar, the partial pressure of O_2 is 207 mbar. It is not as simple as having a pressure of 207 mbar of oxygen in the spacesuit. In the alveoli of your lungs, there is a partial pressure of CO_2 (53 mbar) due to respiration and water vapour (63 mbar) from the lung tissue. This adds up to a total alveolar pressure of 116 mbar. In order to have a partial pressure of oxygen of 207 mbar, the spacesuit must be pressurized by an extra 116 mbar, to 323 mbar O_2. If the spacesuit were only pressurized to 207 mbar with pure oxygen, the partial pressure of O_2 in the lungs would be 207 – 116 = 91 mbar, about 44% of the partial pressure at sea level. This is the same as the partial pressure of oxygen at 5.5 km above sea level, which is 400 m higher than the highest permanently inhabited town in the world, Rinconada, Peru. For an astronaut, working under this low pressure of oxygen would be similar to the experience of high altitude mountain climbing.

▲ Cosmonaut Sergey Volkov dons a spacesuit during a 6 hour and 23 minute spacewalk.

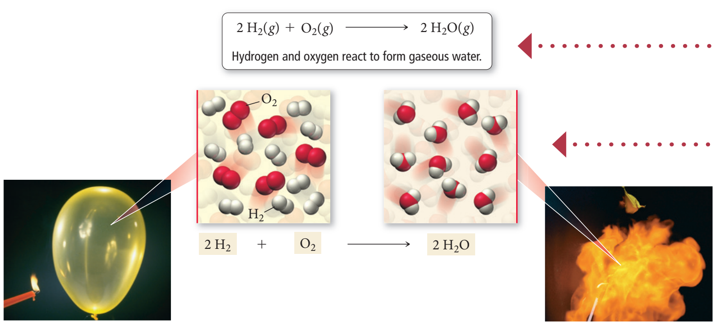
Pioneering artwork makes CONCEPTS CLEAR

Annotated Molecular Art

Many illustrations have three parts:

- a macroscopic image (what you can see with your eyes)
- a molecular image (what the molecules are doing)
- a symbolic representation (how chemists represent the process with symbols and equations)

The goal is for you to connect what you see and experience (the macroscopic world) with the molecules responsible for that world, and with the way chemists represent those molecules. After all, this is what chemistry is all about.



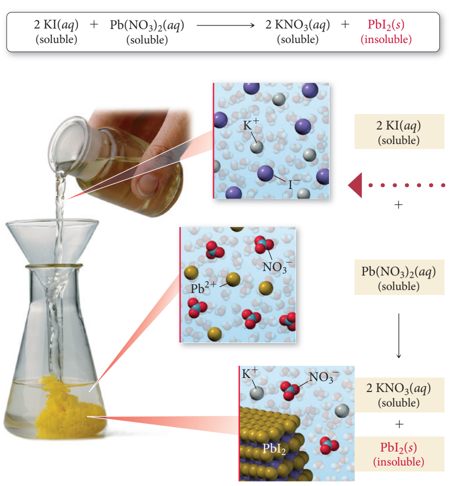
▲ **FIGURE 4.10 Oxidation–Reduction Reaction** The hydrogen in the balloon reacts with oxygen upon ignition to form gaseous water (which is dispersed in the flame).

Symbolic representation

Molecular image

Macroscopic image

▶ **FIGURE 4.7 Precipitation of Lead(II) Iodide** When a potassium iodide solution is mixed with a lead(II) nitrate solution, a yellow lead(II) iodide precipitate forms.



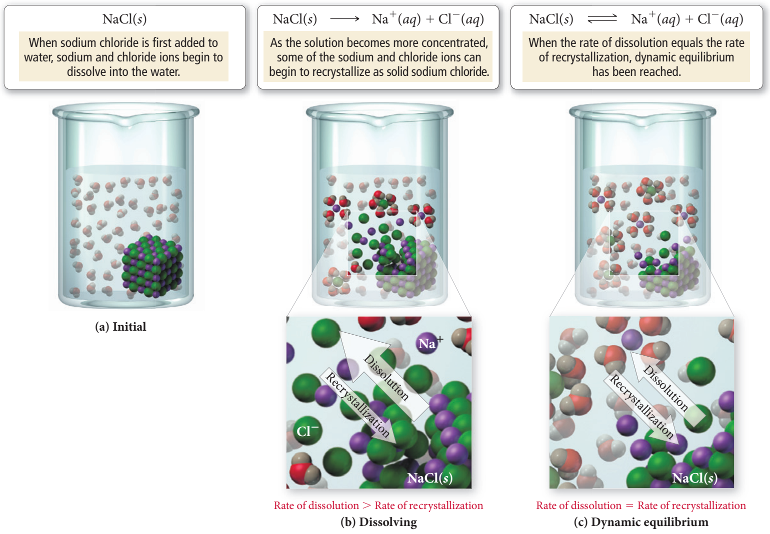
Annotations tell the story of the image concisely.

Molecular image

Macroscopic image

Multipart Images

Multipart images make connections among graphical representations, molecular processes, and the macroscopic world.



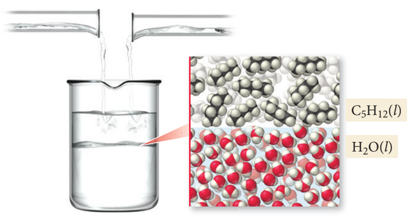
▲ **FIGURE 12.9** **Dissolution of NaCl**

Symbolic representation

Macroscopic image

Molecular image

▶ **FIGURE 11.10** **Polar and Nonpolar Compounds** Water and pentane do not mix because water molecules are polar and pentane molecules are nonpolar.



Graphical representation

Consistent strategies help you SOLVE PROBLEMS

Two-Column Example

A consistent approach to problem solving is used throughout the book.

The molar mass of any element yields the conversion factor between mass (in grams) of that element and the amount (in moles) of that element. For carbon:

$$12.01 \text{ g C} = 1 \text{ mol C} \quad \text{or} \quad \frac{12.01 \text{ g C}}{1 \text{ mol C}} \quad \text{or} \quad \frac{1 \text{ mol C}}{12.01 \text{ g C}}$$

We now have all the tools to count the number of atoms in a sample of an element by weighing it. First, obtain the mass of the sample. Then convert it to the amount in moles using the element's molar mass. Finally, convert to number of atoms using Avogadro's number. The conceptual plan for these kinds of calculations takes the following form:

$$\text{g element} \xrightarrow[\text{of element}]{\text{molar mass}} \text{mol element} \xrightarrow[\text{number}]{\text{Avogadro's}} \text{number of atoms}$$

Example 2.4 demonstrates these conversions.

Notice that numbers with large exponents, such as 6.022×10^{23}, are unbelievably large. Twenty-two copper pennies contain 6.022×10^{23} or 1 mol of copper atoms, but

The **left column** explains how the problem is solved.

The **right column** shows the implementation of the steps explained in the left column.

EXAMPLE 2.4 THE MOLE CONCEPT: CONVERTING FROM MASS TO MOLES AND NUMBER OF ATOMS

Calculate the number of moles of copper atoms and the number of copper atoms that are in 3.10 g of copper.

A **four-part structure ("Sort, Strategize, Solve, Check")** provides you with a framework for analyzing and solving problems.

SORT You are given the mass of copper atoms and asked to find the number of moles of copper atoms and the number of copper atoms.

GIVEN: 3.10 g Cu
FIND: Moles and number of Cu atoms

STRATEGIZE Convert between the mass of an element in grams and the number of moles of atoms of the element with the molar mass.

Then convert from moles to the number of atoms using Avogadro's number.

CONCEPTUAL PLAN

$$\text{g Cu} \xrightarrow{\frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}}} \text{mol Cu} \xrightarrow{\frac{6.022 \times 10^{23} \text{ Cu atoms}}{1 \text{ mol Cu}}} \text{number of Cu atoms}$$

RELATIONSHIPS USED

$63.55 \text{ g Cu} = 1 \text{ mol Cu}$ (Molar mass of copper)
$6.022 \times 10^{23} = 1 \text{ mol}$ (Avogadro's number)

Many problems are solved with a **conceptual plan** that provides a visual outline of the steps leading from the given information to the solution.

SOLVE Follow the conceptual plan to solve the problem. Begin with 3.10 g Cu and multiply by the appropriate conversion factor to obtain the number of moles of copper.

Then multiply the number of moles by Avogadro's number to arrive at the number of copper atoms.

SOLUTION

Number of moles Cu:

$$3.10 \cancel{\text{g Cu}} \times \frac{1 \text{ mol Cu}}{63.55 \cancel{\text{g Cu}}} = 4.88 \times 10^{-2} \text{ mol Cu}$$

Number of Cu atoms:

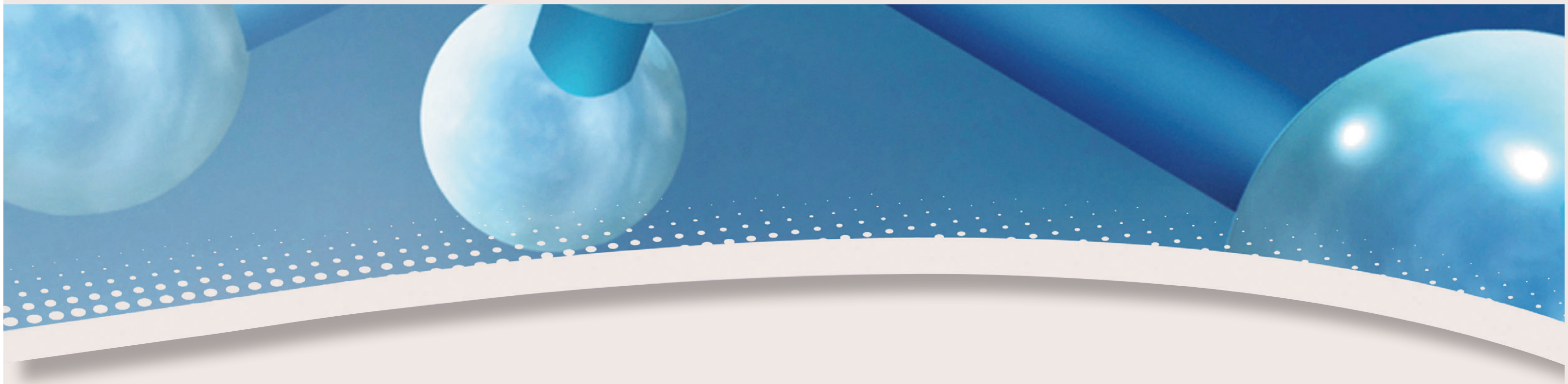
$$4.88 \times 10^{-2} \cancel{\text{mol Cu}} \times \frac{6.022 \times 10^{23} \text{ Cu atoms}}{1 \cancel{\text{mol Cu}}} = 2.94 \times 10^{22} \text{ Cu atoms}$$

CHECK The answer (the number of copper atoms) is less than 6.022×10^{23} (one mole). This is consistent with the given mass of copper atoms, which is less than the molar mass of copper.

FOR PRACTICE 2.4
How many carbon atoms are there in a 1.3 carat diamond? Diamonds are a form of pure carbon. (1 carat = 0.20 grams)

FOR MORE PRACTICE 2.4
Calculate the mass of 2.25×10^{22} tungsten atoms.

Every worked Example is followed by one or more "For Practice" problems that you can try to solve on your own. Answers to "For Practice" problems are in Appendix IV.



Three-Column Example

Problem-Solving Procedure Boxes for important categories of problems enable you to see how the same reasoning applies to different problems.

The formula for the ionic compound composed of *calcium* and chlorine, however, is $CaCl_2$ because Ca always forms 2+ cations and Cl always forms 1− anions in ionic compounds. In order for this compound to be charge-neutral, it must contain one Ca^{2+} cation for every two Cl^- anions.

Summarizing Ionic Compound Formulas

- Ionic compounds always contain positive and negative ions.
- In a chemical formula, the sum of the charges of the positive ions (cations) must equal the sum of the charges of the negative ions (anions).
- A formula reflects the smallest whole-number ratio of ions.

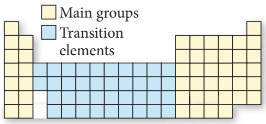
To write the formula for an ionic compound, follow the procedure in the left column in the following example. Two examples of how to apply the procedure are provided in the centre and right columns.

PROCEDURE FOR … Writing Formulas for Ionic Compounds	EXAMPLE 3.2 Writing Formulas for Ionic Compounds Write a formula for the ionic compound that forms between aluminum and oxygen.	EXAMPLE 3.3 Writing Formulas for Ionic Compounds Write a formula for the ionic compound that forms between calcium and oxygen.
1. Write the symbol for the metal cation and its charge followed by the symbol for the nonmetal anion and its charge. Obtain charges from the element's group number in the periodic table (refer to Figure 2.12).	Al^{3+} O^{2-}	Ca^{2+} O^{2-}
2. Adjust the subscript on each cation and anion to balance the overall charge.	Al^{3+} O^{2-} ↓ Al_2O_3	Ca^{2+} O^{2-} ↓ CaO
3. Check that the sum of the charges of the cations equals the sum of the charges of the anions.	cations: 2(3+) = 6+ anions: 3(2−) = 6− The charges balance.	cations: 2+ anions: 2− The charges balance.
	FOR PRACTICE 3.2 Write a formula for the compound formed between potassium and sulfur.	**FOR PRACTICE 3.3** Write a formula for the compound formed between aluminum and nitrogen.

Naming Ionic Compounds The first step in naming an ionic compound is to identify it as one. Ionic compounds are often composed of metals and nonmetals; any time you see a metal and one or more nonmetals together in a chemical formula, assume that you have an ionic compound.

Table 3.2 gives names of common cations and anions. In the case of KBr, the name of the K^+ ion is potassium. For metals that form cations with only one charge, the name of the cation is the same as the metal. For metals that can form cations with different charges, the name of the cation is the name of the metal followed by the charge in roman numerals in brackets. Thus, Fe^{2+} is named *iron(II)* and Fe^{3+} is named *iron(III)*. Many transition metals give ions with different charges (Figure 3.6 ▸). Names for monoatomic anions consist of the **base name** of the element followed by the suffix *–ide*. For example, the base name for bromine is *brom*, and the name of the Br^- ion is *bromide*. The name of KBr is the name of the K^+ cation, followed by the name of the Br^- anion: *potassium bromide*.

The name of the ionic compound is simply the name of the cation followed by the name of the anion.



▲ **FIGURE 3.6 Transition Elements**
Metals that can have different charges in different compounds are usually (but not always) found in the transition elements.

The **general procedure** is shown in the left column.

Two worked examples, side by side, make it easy to see how differences are handled.

Every worked Example is followed by one or more "For Practice" problems that you can try to solve on your own. Answers to "For Practice" problems are in Appendix IV.

End-of-chapter

MATERIAL AIDS STUDY AND TEST PREP

End-of-Chapter Review Section

The end-of-chapter review section helps you study the chapter's concepts and skills in a systematic way that is ideal for test preparation.

Key Concepts

Pressure (5.1, 5.2)

Gas pressure is the force per unit area that results from gas particles colliding with the surfaces around them. Pressure is measured in a number of units, including bar, mbar, mmHg, torr, Pa, psi, and atm.

The Gas Laws (5.3)

The gas laws express relationships between pairs of variables when the other variables are held constant. Boyle's law states that the volume of a gas is inversely proportional to its pressure. Charles's law states that the volume of a gas is directly proportional to its temperature. Avogadro's law states that the volume of a gas is directly proportional to the amount (in moles).

The Ideal Gas Law and Its Applications (5.4, 5.5)

The ideal gas law, $PV = nRT$, gives the relationship among all four gas variables and contains the gas laws within it. We can use the ideal gas law to find one of the four variables given the other three. We can use it to calculate the molar volume of an ideal gas, which is 22.7 L at STP, and to calculate the density and molar mass of a gas.

Mixtures of Gases and Partial Pressures (5.6)

In a mixture of gases, each gas acts independently of the others so that any overall property of the mixture is the sum of the properties of the individual components. The pressure of any individual component is its partial pressure.

Gas Stoichiometry (5.7)

In reactions involving gaseous reactants and products, quantities are often reported in volumes at specified pressures and temperatures. We can convert these quantities to amounts (in moles) using the ideal gas law. Then we can use the stoichiometric coefficients from the balanced equation to determine the stoichiometric amounts of other reactants or products. The general form for these types of calculations is often as follows: volume A $\longrightarrow$ amount A (in moles) $\longrightarrow$ amount B (in moles) $\longrightarrow$ quantity of B (in desired units). In cases where the reaction is carried out at STP, the molar volume at STP (22.7 L $=$ 1 mol) can be used to convert between volume in litres and amount in moles.

Kinetic Molecular Theory and Its Applications (5.8, 5.9)

Kinetic molecular theory is a quantitative model for gases. The theory has three main assumptions: (1) the gas particles are negligibly small, (2) the average kinetic energy of a gas particle is proportional to the temperature in kelvin, and (3) the collision of one gas particle with another is completely elastic (the particles do not stick together). The gas laws all follow from the kinetic molecular theory.

We can also use the theory to derive the expression for the root mean square velocity of gas particles. This velocity is inversely proportional to the molar mass of the gas, and therefore—at a given temperature—smaller gas particles are (on average) moving more quickly than larger ones. The kinetic molecular theory also allows us to predict the mean free path of a gas particle (the distance it travels between collisions) and relative rates of diffusion or effusion.

Real Gases (5.10)

Real gases differ from ideal gases to the extent that they do not always fit the assumptions of kinetic molecular theory. These assumptions tend to break down at high pressures, where the volume is higher than predicted for an ideal gas because the particles are no longer negligibly small compared to the space between them. The assumptions also break down at low temperatures where the pressure is lower than predicted because the attraction between molecules combined with low kinetic energies causes partially inelastic collisions. The van der Waals equation predicts gas properties under nonideal conditions.

▲ The **Key Concepts** section summarizes the chapter's most important ideas.

Key Terms

Section 5.1
pressure (149)

Section 5.2
millimetre of mercury (mmHg) (151)
barometer (151)
torr (151)
pascal (Pa) (151)
atmosphere (atm) (151)
standard pressure (152)
bar (152)
millibar (mbar) (152)
manometer (152)

Section 5.3
Boyle's law (154)
Charles's law (157)
Avogadro's law (159)

Section 5.4
ideal gas law (160)
ideal gas (160)
ideal gas constant (160)

Section 5.5
molar volume (162)
standard temperature and pressure (STP) (162)

Section 5.6
partial pressure (P_n) (166)
Dalton's law of partial pressures (166)
mole fraction (χ_a) (166)
vapour pressure (169)

Section 5.8
kinetic molecular theory (174)

Section 5.9
mean free path (180)
diffusion (180)
effusion (180)
Graham's law of effusion (181)

Section 5.10
van der Waals equation (184)
real gas (184)

▲ **Key Terms** list all of the chapter's boldfaced terms, organized by section in order of appearance, with page references. Definitions are found in the Glossary.

Key Equations and Relationships

Relationship Between Pressure (P), Force (F), and Area (A) (5.2)

$$P = \frac{F}{A}$$

Boyle's Law: Relationship Between Pressure (P) and Volume (V) (5.3)

$$V \propto \frac{1}{P}$$

$$P_1V_1 = P_2V_2$$

Charles's Law: Relationship Between Volume (V) and Temperature (T) (5.3)

$$V \propto T \quad \text{(in K)}$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Avogadro's Law: Relationship Between Volume (V) and Amount in Moles (n) (5.3)

$$V \propto n$$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Ideal Gas Law: Relationship Between Volume (V), Pressure (P), Temperature (T), and Amount (n) (5.4)

$$PV = nRT$$

▲ The **Key Equations and Relationships** section lists each of the key equations and important quantitative relationships from the chapter.

Key Skills

Calculating Internal Energy from Heat and Work (6.3)
• Example 6.1 • For Practice 6.1 • Exercises 41–44, 53–54

Finding Heat from Temperature Changes (6.4)
• Example 6.2 • For Practice 6.2 • For More Practice 6.2 • Exercises 47–48

Thermal Energy Transfer (6.4)
• Example 6.3 • For Practice 6.3 • Exercises 49–50, 65–70

Finding Work from Volume Changes (6.4)
• Example 6.4 • For Practice 6.4 • Exercises 51–52

Finding Pressure–Volume Work for Chemical Reactions Involving Gases (6.4)
• Example 6.5 • For Practice 6.5 • Exercises 53–56

Using Bomb Calorimetry to Calculate $\Delta_r U$ and $\Delta_r H$ (6.5, 6.6)
• Examples 6.6, 6.7 • For Practice 6.6, 6.7 • For More Practice 6.6 • Exercises 73–74

Predicting Endothermic and Exothermic Processes (6.6)
• Example 6.6 • For Practice 6.6 • Exercises 59–60

Determining Heat from ΔH and Stoichiometry (6.6)
• Examples 6.8, 6.12 • For Practice 6.8, 6.12 • For More Practice 6.8 • Exercises 61–64

Finding $\Delta_r H$ Using Calorimetry (6.7)
• Example 6.9 • For Practice 6.9 • Exercises 75–76

Finding $\Delta_r H$ Using Hess's Law (6.8)
• Example 6.10 • For Practice 6.10 • For More Practice 6.10 • Exercises 79–82

▲ The **Key Skills** section lists the major types of problems that you should be able to solve, with the chapter examples that show the techniques needed—along with the "For Practice" problems and end-of-chapter exercises that offer practice in those skills.

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End-of-Chapter Review Exercises

Answers to odd-numbered questions are in Appendix III.

Review Questions

1. Why is molecular geometry important? Give some examples.
2. According to VSEPR theory, what determines the geometry of a molecule?
3. Name and sketch the five basic electron geometries, and state the number of electron groups corresponding to each. What constitutes an *electron group*?
4. Explain the difference between electron geometry and molecular geometry. Under what circumstances are they not the same?
5. Give the correct electron and molecular geometries that correspond to each set of electron groups around the central atom of a molecule:
   a. four electron groups overall; three bonding groups and one lone pair
   c. five electron groups overall; four bonding groups and one lone pair
   d. five electron groups overall; three bonding groups and two lone pairs
   e. five electron groups overall; two bonding groups and three lone pairs
   f. six electron groups overall; five bonding groups and one lone pair
   g. six electron groups overall; four bonding groups and two lone pairs
6. How do you apply VSEPR theory to predict the shape of a molecule with more than one interior atom?
7. How do you determine whether a molecule is polar? Why is

▲ **Review Questions** can be used to review chapter content.

Cumulative Problems

79. Bromine is a highly reactive liquid, while krypton is an inert gas. Explain the difference based on their electron configurations.

80. Potassium is a highly reactive metal, while argon is an inert gas. Explain the difference based on their electron configurations.

81. Suppose you were trying to find a substitute for K^+ in nerve signal transmission. Where would you begin your search? What ions would be most like K^+? For each ion you propose, explain the ways in which it would be similar to K^+ and the ways it would be different. Refer to periodic trends in your discussion.

82. Suppose you were trying to find a substitute for Na^+ in nerve signal transmission. Where would you begin your search? What ions would be most like Na^+? For each ion you propose, explain the ways in which it would be similar to Na^+ and the ways it would be different. Use periodic trends in your discussion.

83. Life on Earth evolved around the element carbon. Based on periodic properties, what two or three elements would you expect to be most like carbon?

▲ **Cumulative Problems** combine material from different parts of the chapter, and often from previous chapters as well, allowing you to see how well you can integrate the course material.

Conceptual Problems

113. Imagine that in another universe, atoms and elements are identical to ours, except that atoms with six valence electrons have particular stability (in contrast to our universe, where atoms with eight valence electrons have particular stability). Give an example of an element in the alternative universe that corresponds to:
   a. a noble gas.
   b. a reactive nonmetal.
   c. a reactive metal.

114. According to Coulomb's law, rank the interactions between the charged particles from lowest potential energy to highest potential energy:
   a. A 1+ charge and a 1− charge separated by 100 pm.
   b. A 2+ charge and a 1− charge separated by 100 pm.
   c. A 1+ charge and a 1+ charge separated by 100 pm.
   d. A 1+ charge and a 1− charge separated by 200 pm.

115. Use the trends in ionization energy and electron affinity to explain why calcium fluoride has the formula CaF_2 and not Ca_2F or CaF.

▲ **Conceptual Problems** let you test your grasp of key chapter concepts, often through reasoning that involves little or no math.

Problems by Topic

Valence Electrons and Dot Structures

33. Write an electron configuration for N. Then write a Lewis structure for N and show which electrons from the electron configuration are included in the Lewis structure.

34. Write an electron configuration for Ne. Then write a Lewis structure for Ne and show which electrons from the electron configuration are included in the Lewis structure.

35. Write a Lewis structure for each atom or ion:
   a. Al **b.** Na^+ **c.** Cl **d.** Cl^-

36. Write a Lewis structure for each atom or ion:
   a. S^{2-} **b.** Mg **c.** Mg^{2+} **d.** P

Ionic Lewis Structures and Lattice Energy

41. Explain the trend in the lattice energies of the alkaline earth metal oxides:

Metal Oxide	Lattice Energy ($kJ\ mol^{-1}$)
MgO	–3795
CaO	–3414
SrO	–3217
BaO	–3029

42. Rubidium iodide has a lattice energy of $-617\ kJ\ mol^{-1}$, while potassium bromide has a lattice energy of $-671\ kJ\ mol^{-1}$. Why is the lattice energy of potassium bromide more exothermic than

▲ **Problems by Topic** are paired, with answers to the odd-numbered questions appearing in Appendix III.

Challenge Problems

97. Consider the densities and atomic radii of the noble gases at 1 bar and 25 °C:

Element	Atomic Radius (pm)	Density ($g\ L^{-1}$)
He	32	0.16
Ne	70	0.81
Ar	98	—
Kr	112	3.38
Xe	130	—
Rn	—	8.96

   a. Estimate the densities of argon and xenon by interpolation from the data.
   b. Provide an estimate of the density of the yet undiscovered element with atomic number 118 by extrapolation from the data.
   c. Use the molar mass of neon to estimate the mass of a neon atom. Then use the atomic radius of neon to calculate the average density of a neon atom. How does this density compare to the density of neon gas? What does this comparison suggest about the nature of neon gas?
   d. Use the densities and molar masses of krypton and neon to calculate the number of atoms of each found in a volume of 1.0 L. Use these values to estimate the number of atoms that occur in 1.0 L of Ar. Now use the molar mass of argon to estimate the density of Ar. How does this estimate compare to that in part (a)?

98. As we have seen, the periodic table is a result of empirical observation (i.e., the periodic law), but quantum theory explains *why* the table is so arranged. Suppose that, in another universe, quantum theory was such that there were one *s* orbital but only two *p* orbitals (instead of three) and only three *d* orbitals (instead of five). Draw out the first four periods of the periodic table in this alternative universe. Which elements would be the equivalent of the noble gases? Halogens? Alkali metals?

99. Consider the metals in the first transition series. Use periodic trends to predict a trend in density as you move to the right across the series.

100. Only trace amounts of the synthetic element darmstadtium, atomic number 110, have been obtained. The element is so highly unstable that no observations of its properties have been possible. Based on its position in the periodic table, propose three different plausible valence electron configurations for this element.

▲ **Challenge Problems** are designed to challenge stronger students.

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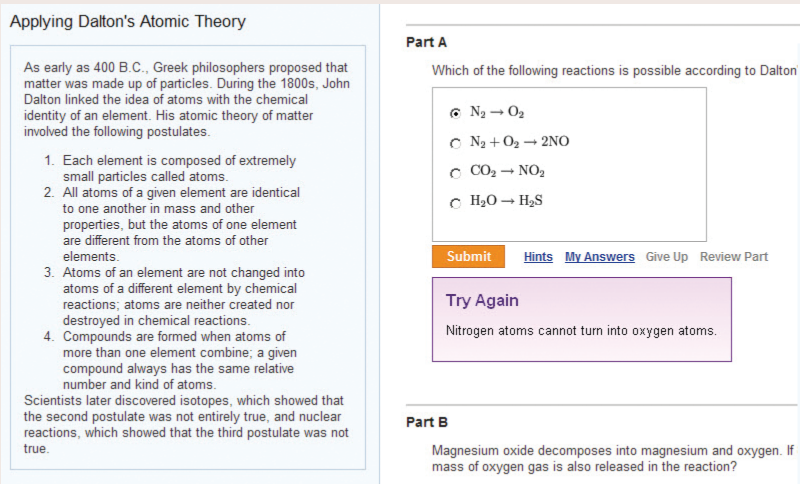

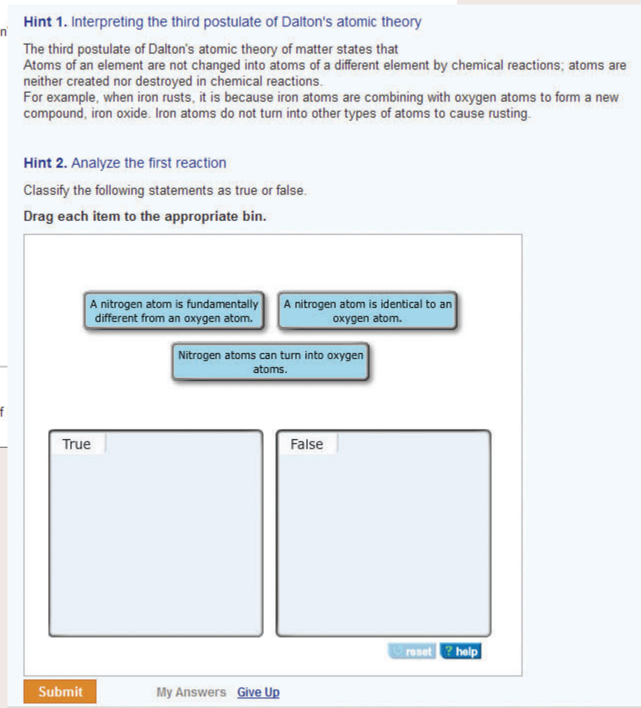

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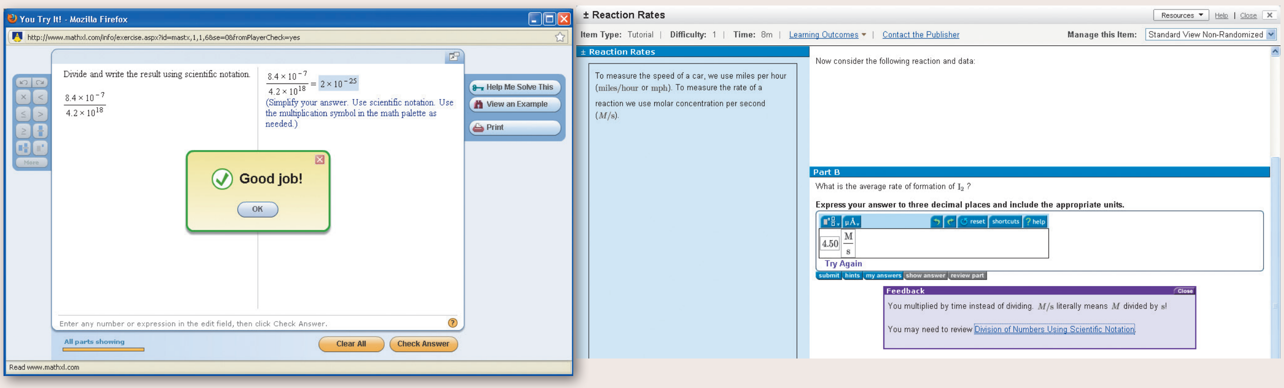

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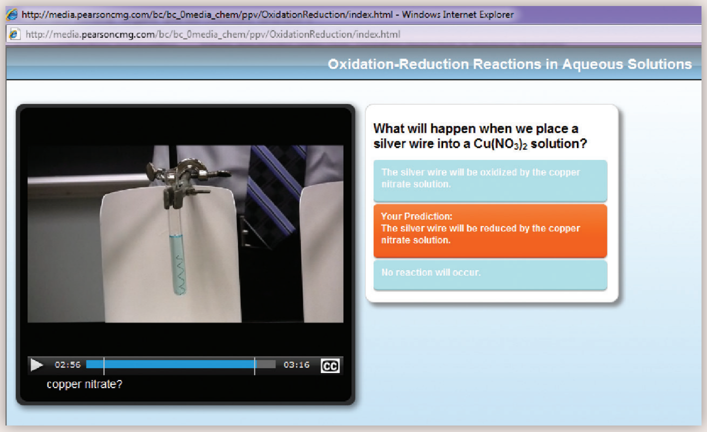

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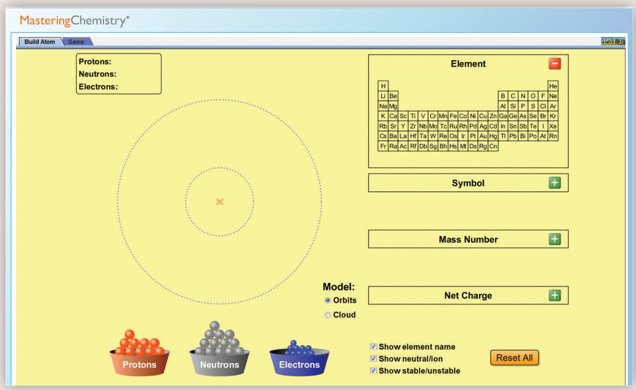

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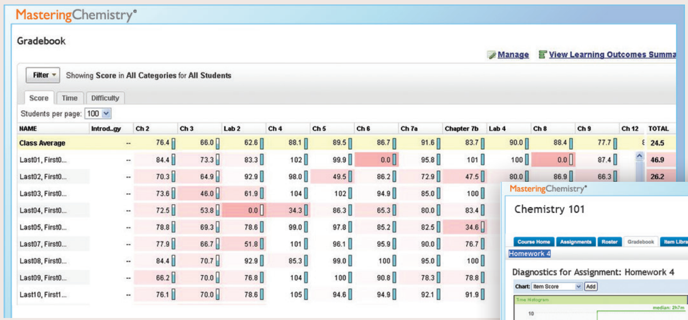

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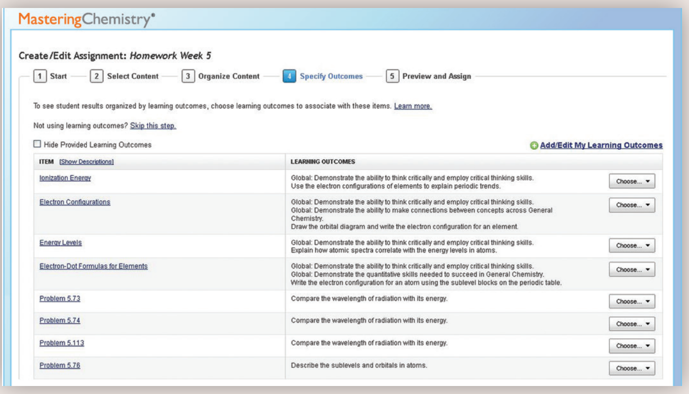

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