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ABSTRACT

This teacher's guide is designed to provide science teachers with the necessary guidance and suggestions for teaching inorganic chemistry. The material in this book can be integrated with the other modules in a sequence that helps students to see that chemistry is a unified science. Contents include: (1) "Periodicity: A Chemical Calendar"; (2) "Structural Chemistry of Metals and Their Compounds"; (3) "Inorganic Molecules"; (4) "Acids and Bases"; (5) "Chemistry of the Transition Elements"; and (6) "Bioinorganic Chemistry". (KHR)

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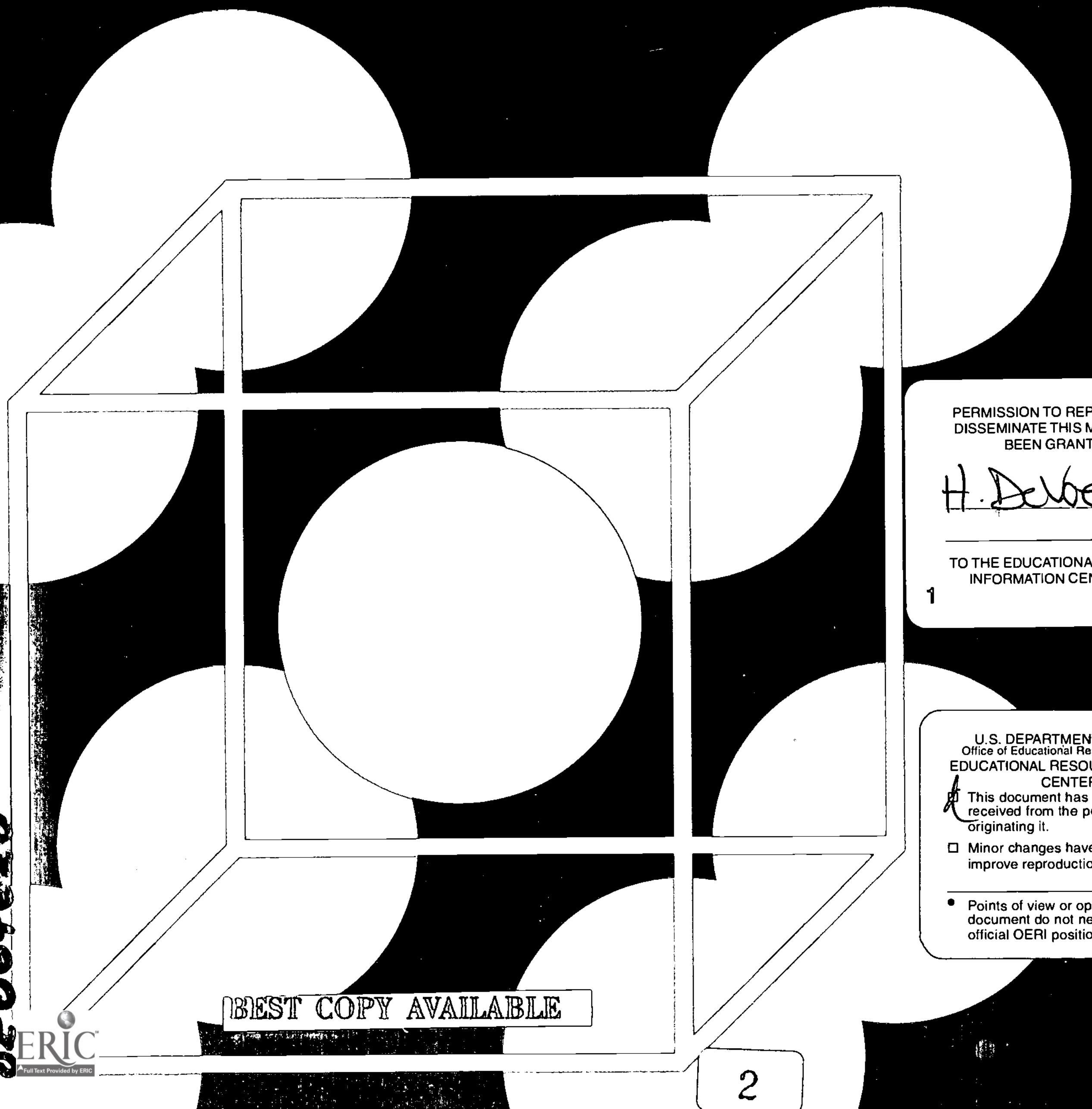
TEACHER'S GUIDE

DIVERSITY AND PERIODICITY

AN INORGANIC CHEMISTRY MODULE



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TEACHER'S GUIDE

DIVERSITY AND PERIODICITY

AN INORGANIC CHEMISTRY MODULE

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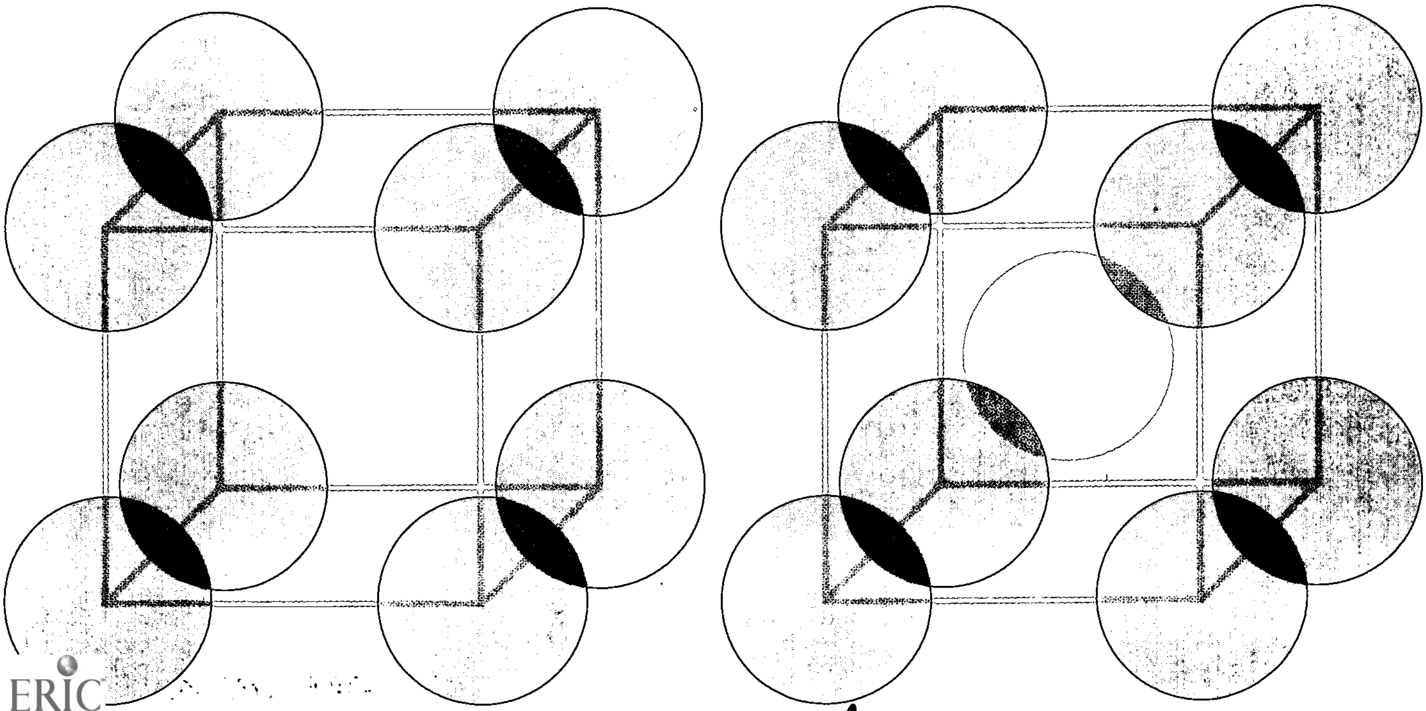

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TEACHER'S GUIDE

DIVERSITY AND PERIODICITY:
AN INORGANIC CHEMISTRY MODULE

JAMES HUHEEY

When not driving his red and white truck up and down the mountains of western North Carolina in search of "the perfect salamander," ratchet-jawing with the 18-wheelers out on "that ol' boulevard" with his CB, or pounding a typewriter to keep his editor happy, Jim Huheey finds time to do some of the other things inorganic chemists do: research on electronegativity and related topics, lecture preparations, faculty meetings, Sunday afternoon football games, etc., etc. Ever since he joined the University of Maryland from the University of Illinois via Worcester Polytech, Jim Huheey has been a major contributor to the mainstream of the department life. Characterizing himself (in a friendly needle to the organic chemists) as a chemist of *the other* 105 elements, his work in the past has been geared mostly to the college and university levels. Lately he has extended his range to encompass high school as well.

Capping his many achievements, in 1971 he received the Young Chemistry Teachers' Award sponsored by the District of Columbia division of the American Institute of Chemists; and in 1972 his publishers brought out his second book, a college text titled *Inorganic Chemistry: Principles of Structure and Reactivity*.

AMADO SANDOVAL

Since 1969 when he went to the University of Maryland as a Chemistry Teaching Associate, Amado (Pepe) Sandoval has been involved with chemical education in the Maryland-D.C. area: first as a teaching associate, then as a student-teacher supervisor, teacher workshop instructor, and now as a teacher at Wilde Lake and Centennial High Schools in Howard Co., Maryland. While at Wilde Lake High School he has been involved in the development of the IAC program where he developed teacher guide material and independent study guides for all IAC modules.

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Introducing Diversity and Periodicity

Diversity and Periodicity: An Inorganic Chemistry Module illustrates one of the many facets of modern chemistry. As with all of the other IAC modules, it is written to stand alone, requiring only the prior coverage of *Reactions and Reason: An Introductory Chemistry Module*. You will also find that the material here can be integrated with the other modules in a sequence that helps the student to see that chemistry is a unified, though multifaceted, science.

When should this module be taught? Some IAC teachers use it early with good results, and others work it in equally satisfactorily after teaching several other modules. We leave the timing to you. We would emphasize, however, that the old notion that inorganic chemistry is "basic" or "simple," and that therefore it should be taught early, is spurious. The idea stems from the day when the first course in college chemistry was called "inorganic chemistry" and consisted of memorizing some reactions of various elements. As the IAC team has tried to show in all of these modules, there are basic and fundamental concepts in every branch of chemistry, and there are very sophisticated, even philosophical, concepts to be derived from all areas.

As we worked on *Diversity and Periodicity*, in addition to the desire to make the material interesting and attractive, we found ourselves asking the questions: What do inorganic chemists do? How do they go about their work? How does their work impinge upon the society in which they live? We have tried to answer these questions for the student, for we feel that an understanding of what a chemist does and thinks is as important a part of the student's education as knowing the product of a particular reaction. In fact, long after most of your students have lost active contact with chemistry (only a very small fraction will become chemists), they will be well-informed and clear-thinking citizens making decisions on this nation's priorities, many of them dealing with science.

This module covers old favorites of inorganic chemistry, including discussion of ammonia and sulfur trioxide and the titration of acids and bases. In addition, there are syntheses that may be new to you in a high-school context, such as the preparation of coordination compounds, a very important class of inorganic compounds. You may also be surprised to be discussing molecules such as hemoglobin in the context of inorganic chemistry. Even old war-horses—NH_3 and SO_3, for instance—are considered in light of current and related topics, such as nitrogen-fixation (by plants and people) and air pollution, as well as more basic concepts. We have not broadened our scope in this way to prove how clever we are; all of these topics—coordination compounds, hemoglobin, nitrogen-fixation, and air pollution—are of paramount concern to inorganic chemistry today.

We have tried to provide you with the necessary guidance and suggestions for teaching this module. We hope you will enjoy teaching *Diversity and Periodicity: An Inorganic Chemistry Module* and that your students will enjoy working with it.

Special Features in the Student Module

Metric System Le Systeme Internationale (SI) is used throughout the IAC program. As you work with this module, you may wish to review some points of the metric system as presented in *Reactions and Reason: An Introductory Chemistry Module* (see section *A-8* and Appendix II). There is a metric units chart in the appendix of the student module that students can easily refer to.

Time Machine A feature we call the *Time Machine* appears in the IAC modules in order to show chemistry in a broader context. For some students, this may provide a handle on particular aspects of chemistry by establishing the social-cultural-political framework in which significant progress was made in chemistry. Students may enjoy suggesting other events in chemistry around which to create *Time Machines* of their own.

Cartoons A popular feature of the IAC program is the use of chemistry cartoons. These cartoons give students a chance to remember specific points of chemistry in another important way—by humor. Suggest that your students create other chemistry cartoons for their classmates to enjoy.

Safety Laboratory safety is a special concern in any chemistry course. In addition to including safety discussion and guidelines in the appendix of each student module and teacher's guide experiments have been developed in a way to eliminate potentially dangerous chemicals or procedures. Moreover, each experiment that might present a hazard—through fumes, corrosive chemicals, use of a flame, or other conditions—has been marked with a safety symbol to alert students and teachers to use added, reasonable caution. Caution statements, in bold type, also appear in experiments to specifically instruct the student on the care required.

Selected Readings Articles and books that tie in with the topics discussed in the IAC program have been listed in the appendix of the student module as well as in the teacher's guide. Encourage your students to use this section. You may wish to suggest other material that you yourself have found interesting and enjoyable.

Illustrations and Photographs The module is extensively illustrated to provide relevant and stimulating visual material to enable students to relate chemistry to everyday life, as well as to provide material for provocative discussion. In using some of these illustrations, it is not the intention of IAC to endorse any particular product or brand, but only to relate chemistry to life outside the classroom. As you proceed through each section, encourage students to collect, display, and discuss photos and illustrations that provoke further discussion.

Questions A number of questions have been interspersed throughout the student module, in addition to the questions that are naturally built into the narratives and the laboratory experiments. You will find some of these in specifically marked sections in the student module. These questions can be used in a variety of ways as you see fit. They are not planned as tests—remember, the IAC program is designed so that mastery of the content and skills can be achieved through the repeated reinforcement of ideas and procedures encountered by students as they progress through the various modules.

Managing the Laboratory

In the teacher's guide, hints and suggestions are given for managing each experiment in the laboratory. Share as many of these hints as possible with your students. Allow them to participate fully in successful laboratory management. Make sure that you rotate assignments so that all students get a chance to experience this type of participation.

Preparations and Supplies Student aides can be helpful in preparing solutions, labeling and filling bottles, cleaning glassware, and testing experiments. (You should still test each experiment in the module to determine if any revision is needed to meet the needs of your students.)

Cleaning Up Involve your students in putting away equipment, washing up glassware, and storing material for the next time it is to be used.

Taking care of equipment is part of responsibilities we seek to foster in the students' outside environment.

Laboratory Reports You may have your own methods of student reporting. We are including some of the suggestions that IAC teachers have found successful in the past. It is helpful for students to keep a laboratory notebook. A quadrille-ruled laboratory notebook with a sheet of carbon paper allows a student to produce two data sheets and report summary copies. One copy of each page can be permanently retained in the notebook, while the duplicate copy can be submitted for evaluation or tabulation.

A realistic view of laboratory work suggests that, in the most fundamental sense, there are no "wrong" laboratory results. All students obtain results consistent with particular experimental conditions (either correct or incorrect) that they establish. Careful work will yield more precise results, of course. Encourage each student to take personal pride in experimental work. If students disagree on a result, discuss the factors that might account for the difference. A student who provides a thoughtful analysis of why a particular result turned out to be "different" (incomplete drying, a portion of the original sample was spilled, etc.) deserves credit for such interpretation.

Laboratory Safety To use the IAC program safely, you should become thoroughly familiar with all student activities in the laboratory. Do all the experiments and carry out all the demonstrations yourself before presenting them to your class. We have tested each experiment and have suggested the use of chemicals that present the least chance of a problem in laboratory safety. This teacher's guide has many suggestions for helping you provide your students with safe laboratory experiences. Have the students read the safety section in the appendix of the student module. Conduct a brief review of laboratory safety before allowing them to encounter their first laboratory experience in this module. Review safety procedures when necessary and discuss caution and safety each time a safety symbol appears in the student text.

Materials for IAC In light of increasing costs for equipment and supplies, as well as decreasing school budgets, we have tried to produce a materials list that reflects only the quantities needed to do the experiments, with minimal surplus. Thus, the laboratory preparation sections contain instructions for only a 10-20 percent surplus of reagents. Add enough materials for student repeats and preparations errors.

Evaluating Student Performance

There are many ways of evaluating your students' performances. One of the most important forms of evaluation is observing your students as they proceed through the IAC program. IAC has developed skill tests and knowledge tests for use with this module. These test items have been suggested and tested by IAC classroom chemistry teachers. You are encouraged to add these to your own means of student evaluation.

In addition to the questions incorporated in the student module text and illustration captions, there are suggestions for evaluation at the end of each module section in the teacher's guide. The module tests are at the end of the teacher's guide. Answers to all of the evaluation items are included to help you in your classroom discusion and evaluation.

Module Concepts

INORGANIC CHEMISTRY

- Inorganic chemistry is a study of the properties of all the elements.
- Each element is unique, yet it is possible to provide a unifying description of all elements.

PERIODICITY: A CHEMICAL CALENDAR

- Many properties of elements recur in fixed cycles; these are called *periodic properties.*
- The periodic chart displays the elements arranged by their increasing atomic number.
- Elements may be divided into two groups—metals and nonmetals.
- Knowledge of periodic properties enables chemists to make predictions about the physical and chemical properties of elements.
- All elements occurring in a single column of the periodic table have similar properties; they form a family of elements.
- Each family of elements differs from the others in the tendency to gain or lose electrons; this is recognized in the combining ability of these elements.

STRUCTURAL CHEMISTRY OF METALS AND THEIR COMPOUNDS

- The physical properties of metals are related to the bonding between atoms and the lattice structure of the solid.
- Different lattice structures may have a greater number of atoms of the same size that could be packed in a given volume.
- The physical properties of ionic compounds are related to the bonding between ions and the lattice structure of the solid.
- Similar ionic compounds may have different lattice structures.

INORGANIC MOLECULES

- Inorganic molecules can be formed by the covalent bonding between nonmetallic elements.
- Physical and chemical properties of molecules are related to their geometry and structure; these may be influenced by nonbonding electron pairs.
- The partial electronic charges upon a polar molecule may attract other polar molecules and form a "bond," such as the hydrogen bond.
- The synthesis of some nonmetallic compounds would not be feasible without the use of catalysts.

ACIDS AND BASES

- The properties of an acid can be looked upon as being opposite to the properties of a base.
- One definition of an acid is a compound that donates hydrogen ions to water in solution to form hydronium ions, H_3O^+.
- A base is an acceptor of hydrogen ions; basic solutions result when the solute removes hydrogen ions from water molecules, leaving hydroxide ions behind.
- Acids and bases react with each other to neutralize the properties of both; water and salt are neutralization products.
- Acid-base neutralization reactions are quantitative.
- Acids and bases are present in many household cleaners because of their powerful chemical action.
- A Lewis acid is an electron pair acceptor and a Lewis base an electron pair donor.

CHEMISTRY OF THE TRANSITION ELEMENTS

- Transition elements occupy a specific area of the periodic table that lies between the alkali and alkaline earth metals (Groups IA and IIA) and the remaining representative metals and nonmetals (Groups IIIA–VIIIA).
- The transition metals tend to lose electrons, forming positive ions, and in doing so are *oxidized.*
- Transition or other metallic ions may gain electrons to become metal atoms and in doing so are *reduced.*
- Transition metals may be oxidized or reduced in more than one step; that is, they usually have more than one stable *oxidation state.*
- When accepting electrons, transition metal ions act as *oxidizing agents;* when giving up electrons, they act as *reducing agents.*
- A transition metal ion may function as a Lewis acid, accept electron pairs from a ligand, and form complexes (coordination compounds) by covalent bonding.
- Molecules or anions able to donate an electron pair (a Lewis base) may act as ligands.
- Coordination compounds often display characteristic properties, such as colors and geometry.

- Metallic ions may be removed from further chemical reaction by formation of coordination complexes.
- Solvent extraction may be used to effect the separation of metals. A desired metal reacts with a particular ligand to form a complex that has a greater solubility in one solvent than another.
- The distinct color of coordination compounds may be used as a specific test for the presence of a given chemical species.

BIOINORGANIC CHEMISTRY

- Bioinorganic chemistry is concerned with the role of inorganic chemicals in biological systems.
- The majority of environmental pollutants are inorganic in nature.
- There are about 20 elements essential to living organisms; some of these are needed in trace amounts.
- Metalloenzymes contain specific metal ions. These may be displaced by other metal ions that "poison" the enzyme.
- The oxygen transport to tissues of higher animals involves a coordination compound of the ligand oxygen (O_2) and the iron ion (Fe^{2+}) in the heme group.
- The driving force of living systems is the oxidation of carbohydrates and similar substances. Many redox reactions that occur in the electron transport system are made possible by enzymes that contain metallic ions.

A SUMMING UP

- Alchemy and the processing of metallic ores provided the "roots" for inorganic chemistry. Inorganic chemistry overlaps other areas of chemistry, such as analytical, physical, and organic chemistry, and recently has extended into new areas, such as bioinorganic chemistry.

Module Objectives

We have attempted to group module objectives in three broad categories: concept-centered, attitude-centered, and skill-centered. The categories are not mutually exclusive; there is considerable overlap. The conditions for accomplishing each objective are not given, since they can easily be found in the respective section in the module. Note also that concept and skill objectives are more specific than those in the affective domain. It is very difficult to classify objectives in this way, but we have been encouraged to do so by classroom teachers, who have helped in this difficult task.

The objectives identified here should provide you with a useful starting point in clarifying your own goals in teaching this module. We encourage you to identify alternative objectives, using this list as a point of departure. Assessment items can be found at the end of major sections in the student module in the form of *Problems.* Other *Evaluation Items* are included after each major section of this guide and in the form of module tests for knowledge and skill objectives located in the teacher's guide appendix.

Concept-Centered Objectives	*Attitude-Centered Objectives*	*Skill-Centered Objectives*
	INTRODUCING INORGANIC CHEMISTRY	
● Give a definition of inorganic chemistry and relate this to the other areas of chemistry.	● Realize that the study of the properties of elements enables the chemist to group them in a manner that might predict the properties and existence of other elements that are not yet discovered.	
	PERIODICITY: A CHEMICAL CALENDAR	
I-1 ● Give examples of cyclic events that "predict" weekly or monthly behavior. **I-2** ● Identify the characteristic properties of metals and nonmetals. ● Distinguish between metals and nonmetals on the basis of their physical properties. ● Give examples of metals, nonmetals, and metalloids that are useful because of malleability, ductility, or absence of these characteristics. **I-3** ● Describe the increase or decrease in ionization energy as a periodic function.	● Sense the beauty of the order of the universe as exemplified by the periodic law. ● Recognize the usefulness of being able to predict properties of chemical elements. ● Be aware that nature does not always allow us to "pigeonhole" things, whether they be elements or people. ● Enjoy the humor to be found in otherwise serious material by using cartoons and puns. ● Discuss the "amazing accuracy" of Mendeleev's predictions as well as their importance to the field of chemistry. ● Appreciate the usefulness of having a unifying theme such as the periodic law in inorganic chemistry or any science.	**I-4** ● Graph atomic number versus ionization energy data. **I-6** ● Outline and identify the main sections of a standard periodic table. **I-9** ● Determine relative chemical reactivity of the halogens through redox reactions.

Concept-Centered Objectives	*Attitude-Centered Objectives*	*Skill-Centered Objectives*
I-4 ● Determine the relation between ionization energy and atomic number in a chemical family and a chemical period. ● Point out the trends of ionization energy changes within different periods and groups of elements.		
I-7 ● Provide an explanation for the increase in length of the repeating cycle with the addition of the transition elements.		
I-8 ● Predict an element's relative chemical reactivity based on the position of the element in the periodic table.		
I-10 ● Use the relationships in the periodic table to predict an element's valence. ● Use an element's valence to write formulas of compounds.		

STRUCTURAL CHEMISTRY OF METALS AND THEIR COMPOUNDS

I-11 ● Explain the malleability of metals as a function of their metallic bonding. **I-13** ● Define coordination number and efficiency of packing, and explain how they are related. **I-14** ● Determine the coordination number of metallic lattice structures. **I-15** ● Recognize the relation between the properties of metals and the arrangement of their atoms and electrons. ● Explain the malleability and ductility of metals by the ease of "layer glide" in the crystal system.	● Recognize that different materials may have different properties depending upon their structure and bonding. ● Appreciate the symmetry and regularity of the various crystal structures. ● Think of the difficulty involved in magnifying atoms *five million* times. ● Recognize the relationship between the properties of ionic compounds and their structures.	**I-12** ● Identify and test physical properties of metals and nonmetals. ● Classify metals and nonmetals on the basis of their physical properties. **I-14** ● Construct three-dimensional models of metallic lattices. ● Calculate the packing efficiency of a metallic lattice structure.

Concept-Centered Objectives	Attitude-Centered Objectives	Skill-Centered Objectives
I-17 ● Qualitatively describe ionic bonding. (Note: Ongoing concepts and objectives are continuously developed through sections *I-18, I-19,* and *I-20.*) **I-17, 18, 19** ● Determine the coordination number of ionic lattice structures. **I-17, 18, 19, 20** ● Determine the common physical properties of ionic compounds. ● Compare physical properties of metals and nonmetals with those of ionic compounds. ● Explain physical properties of ionic crystals in terms of their ionic lattices.		**I-16, 17** ● Build models of ionic packing systems. ● "Grow" ionic crystals. **I-16, 17, 18** ● Construct three-dimensional models of ionic lattices. ● Calculate the packing efficiency of ionic lattice structures.

INORGANIC MOLECULES

I-21 ● Identify covalent bonds as those that result from the sharing of electrons. ● Recognize the existence of molecules that do not conform to the octet rule. **I-23** ● Predict shapes of molecules, including those with nonbonding pairs of electrons. ● Determine correct formulas of some compounds. **I-24** ● Relate the physical properties of molecules to their molecular geometry. **I-25** ● List common important inorganic molecules and indicate why they are important.	● Understand that the chemist can describe things in many ways, from very sophisticated mathematical models to simple, everyday words. As long as no misunderstanding occurs, all of the methods are useful. ● Recognize the reasons that model building is important. ● Realize and discuss implications of changing the geometry, and hence the properties, of essential life-supporting molecules. ● Appreciate the beauty of snowflakes and the relationship of their symmetry to molecular properties. ● Realize that whether a chemical compound is "good" or "bad" depends upon where we find it.	**I-23** ● Write line formulas for molecules. ● Write electron-dot structures for molecules. ● Construct ball-and-stick models of molecules. ● Use magnets as models to determine molecular geometry. ● Use balloons as models to determine molecular geometry.

Concept-Centered Objectives	*Attitude-Centered Objectives*	*Skill-Centered Objectives*
	ACIDS AND BASES	
I-26 ● Define an acid and write an equation for the dissociation of an acid in H_2O. ● Define a base and write an equation for the dissociation of a base in H_2O. ● Write an equation illustrating a neutralization reaction. ● Explain the role of an indicator in the titration process. ● Give examples of acids and bases. **I-28** ● Explain the role of acids and bases in household cleaners. **I-29** ● Write an acid-base reaction using electron-dot formulas to illustrate the Lewis principle.	● Recognize the importance and numerous uses of acids and bases in everyday life. ● Think of the importance of the analysis of commercial products, environmental pollutants, and similar important chemicals in your life.	**I-27** ● Perform an acid-base titration to a correct indicator end point. ● Calculate the percent acid or base in household cleaners. **I-29** ● State the Lewis definition for an acid and a base. ● Give examples of Lewis acids and bases.
	CHEMISTRY OF THE TRANSITION ELEMENTS	
I-30 ● Define oxidation and reduction in terms of electron transfer (gain and loss). ● Determine the oxidation state of elements in compounds and ions. **I-31** ● Identify the two components that make up a coordination compound or complex. ● Define a ligand and give two examples. **I-33** ● List the ligand/metal ratio in common coordination compounds. ● Relate geometry of a coordination complex to its ligand/metal ratio. **I-34** ● State a useful application of coordination compounds.	● Appreciate the beautiful colors of some coordination compounds. ● Recognize the increasing use of transition elements in our modern-day society ● Appreciate the structural beauty of elements and compounds as viewed under the electron microscope.	**I-30** ● Write equations to illustrate the oxidation and reduction of elements. **I-32** ● Synthesize coordination compounds. **I-35** ● Determine the effect of different metal ions on the sudsing ability of soap solutions.

Concept-Centered Objectives	*Attitude-Centered Objectives*	*Skill-Centered Objectives*
I-37 ● Explain how "solvent extraction" is used to separate two metallic coordination compounds.		**I-36** ● Determine the effectiveness of different chemicals as water softeners.
I-38 ● Suggest a method to separate a mixture of coordination compounds using the technique of solvent extraction.		**I-38** ● Use solvent extraction to separate a mixture of compounds.
I-39 ● State a useful analytical application of coordination compounds.		**I-40** ● Test qualitatively for the presence or absence of lead in a sample.

BIOINORGANIC CHEMISTRY

I-41 ● Give examples or write equations to show how inorganic materials can act as environmental pollutants.	● Recognize that many pollutants are inorganic materials. ● Realize the importance of inorganic materials to survival of the living organism.	
I-43 ● Explain the function of metallic elements in biochemical enzymes.	● Recognize that knowledge the chemist obtains in the laboratory can help us understand processes in living systems, including our own bodies.	
I-44 ● Explain the function of hemoglobin and myoglobin in respiration. ● Explain the ligand-complexing role of iron in these molecules.	● Think of your body as a beautifully complex system of chemical reactions integrated like a giant chemical factory.	
I-45 ● Give an example of the role of redox reactions in life processes. ● Illustrate by equation the redox behavior of Fe in the cytochrome c system.		

A SUMMING UP

● Give examples of the elements stating how they form complexes that have important functions in everyday life.	● Realize the valuable role that some knowledge of inorganic chemistry can play in your life.	

Teaching Diversity and Periodicity

Studying inorganic chemistry is a serious business—the study of the *other 106 elements,* as we have noted elsewhere—yet it offers many opportunities for humor and outright laughter. For better or worse, this module is rich in cartoons and puns. We have included some cartoons of our own, some suggested by IAC students and teachers, and others through the courtesy of professional cartoonists. If they provide a bit of a break from the rigors of learning chemistry, perhaps students will be more inclined to try again with something they don't understand. As confirmed punsters, we feel that the worse the puns are, the better. More seriously, it is possible that some of the puns may serve to stimulate students' curiosity or to help them remember useful information.

Introducing Inorganic Chemistry

The student first encounters inorganic chemistry in this module in the person of Dmitri Mendeleev and his periodic tables. Mendeleev's predictions were made more than a century ago. Students might like to speculate on how long it took Mendeleev to develop his tables—this might be cut short by looking into an encyclopedia or a history of chemistry—and what the use of a computer or even a calculator might have meant to him. We will look more closely at Mendeleev, beginning in section *I-3*.

The fundamental themes with which Mendeleev worked and which are common to all of inorganic chemistry are summed up in these two ideas, presented in the opening section:

1. Inorganic chemistry is a study of the properties of all the elements.
2. Each element is unique, yet it is possible to provide a unifying description of all the elements.

You may wish to prepare an overhead transparency which duplicates the versions of Mendeleev's first periodic table that appear on page 2 of the student module. This transparency can be used in discussions of the introductory section and the first section, *Periodicity: A Chemical Calendar*.

Periodicity: A Chemical Calendar

This section develops ideas about the periodic properties of the elements. References to cyclic phenomena and everyday events help to clarify periodicity. Elements are classified first as metals and nonmetals and later as families on the basis of recurring physical and chemical properties. Experiments reinforce family relationships. Eventually the students move on to a miniexperiment in their investigation of periodic properties, especially ionization energies.

Mendeleev's contribution of the periodic table is complemented by a discussion of his successful predictions of elements that were yet to be discovered. The transition metals, as well as the position of hydrogen in the periodic table, are discussed with a reference to ionization energies. The concluding section considers the reactivity of the elements in terms of their positions in the periodic table, valence electrons, and some simple types of compounds that may be formed.

A subtle emphasis on the importance of the transition elements begins in the later parts of this section. Throughout the module the transition elements continue to surface as a most significant group of elements. The transition elements enter into an investigation of enzymes, Lewis acids, coordination compounds, the structures of metals, and, finally, the life processes.

I-1 CHEMICAL CYCLES

Although the weekly cycle and the calendar are valid counterparts of the periodicity of elements and the periodic chart, the analogy in the module discussion may be limited by the fact that weekly cycles are somewhat less well defined than they used to be. Different areas of our culture use different cycles.

The introduction to the periodic chart here is closely related to the film *Chemical Families*. The film is a CHEM Study film, available for rental or purchase from Modern Learning Aids.* Show this film as an introduction to periodicity. It illustrates experiments that are nearly impossible to reproduce in the high-school laboratory and provides an interesting lead-in to periodicity.

I-2 ULTRAMAN VS. FLINTSTONES

In the discussion of the properties of metals and nonmetals (conductivity, luster, malleability, ductility) use actual samples to demonstrate these properties. For instance, hammer a piece of copper wire or foil or cut a piece of sodium.

The conductivity of heat by metals can be illustrated by the fact that metals feel cool at room temperature. They are conducting heat away from the hand. On the other hand (pun intended!), when hot they feel even hotter for the same reason. Have you ever touched a hot pot lid instead of the insulating knob?

If time permits, a discussion of properties of metals can be carried further. Most metals are solids, the exception being mercury (gallium and cesium have low melting points, too, 29.8°C and 28.5°C, respectively). Most metals are rather dense, ranging from 0.53 g/cm^3 for lithium to 22.5 g/cm^3 for osmium. Prompt your students to talk about metals with which they are familiar. Have them discuss the use of these metals in everyday life.

The nonmetallic elements are less likely to be observed in everyday activities. Many of their useful properties also happen to be exhibited by some compounds. For example, sulfur could be melted, cast in molds, and used as a nonconducting insulator on high-voltage electrical lines, but so can glass, and the latter is stronger.

I-3 PERIODIC PROPERTIES

In order to illustrate periodicity, properties that are familiar to the student who has not studied much chemistry were selected. However, the "best" properties to illustrate periodicity are to be found in the properties of the isolated atoms. Ionization energy is used in the present discussion, although it is less intuitively related to the student's prior experience than the other properties.

Allow the students to investigate whether other physical properties of elements are periodic in nature or not. The elemental values for these properties (melting point, density, etc.) may be found by the student in handbooks of chemistry and used in graphs similar to those on pages 10 and 11 of the student module.

MINIEXPERIMENT
I-4 PERIODICITY

This miniexperiment permits students to recognize a periodic property through plotting ionization energy data.

Concepts In doing this miniexperiment, a student will encounter these important ideas:*

- Ionization energy is a periodic property of chemical elements.
- The basic repeating cycle of periodicity is that of eight elements.

Objectives After completing this miniexperiment, a student should be able to:*

- Define ionization energy.
- Graph ionization energy versus atomic number data.
- Determine the relationship between atomic number and ionization energy in a chemical family and a chemical period.

Estimated Time One period (45 minutes), including the discussion

Student Grouping Individuals

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*Modern Learning Aids Division, Ward's Natural Science Establishment, Inc. P.O. Box 1712, Rochester, New York 14603.

*This statement appears only with this first miniexperiment, but it applies each time this section appears in an experiment or miniexperiment, unless otherwise noted.

Materials You will need the following materials for a class of thirty students:*

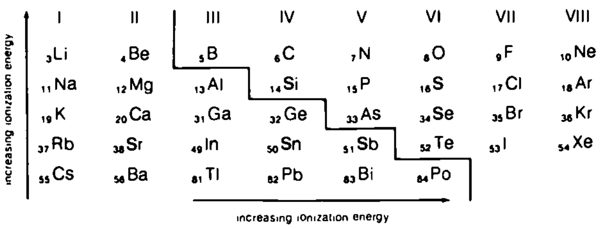
- 30 sheets of linear graph paper
- 15 pairs scissors
- 1 roll of transparent tape (or glue)

Advance Preparation See *Laboratory Tips.*

Prelab Discussion The discussion of periodic trends in sections *I-2* and *I-3* leads naturally into this miniexperiment. Prepare a transparency for an overhead projector that shows the graph axes. Or you might simply put the graph axes on the chalkboard. Students may benefit by reviewing section *A-38* of *Reactions and Reason: An Introductory Chemistry Module.*

Laboratory Tips The use of only the representative elements presents some problems, since it requires that the atomic numbers jump from 10 to 31, 38 to 49, etc. (Note that the number 31 in the student's miniexperiment instructions is not a typographical error!) A reference to only the representative elements also makes it necessary to postpone any discussion of more than half the chemical elements. The alternative is to consider all the elements immediately, which leads to unnecessary complications. The addition of the transition metals and of the longer periods is discussed in section *I-7.*

Range of Results When the elements are arranged according to the directions in the student module, the following results can be seen: (1) elements with low ionization energies, such as sodium and potassium, are found to the left of each cycle; and (2) the elements with high ionization energies, such as fluorine and chlorine, are found to the right of each cycle. The logical arrangement of elements based on this information forms a primitive sort of periodic table.



I	II	III	IV	V	VI	VII	VIII
$_{3}$Li	$_{4}$Be	$_{5}$B	$_{6}$C	$_{7}$N	$_{8}$O	$_{9}$F	$_{10}$Ne
$_{11}$Na	$_{12}$Mg	$_{13}$Al	$_{14}$Si	$_{15}$P	$_{16}$S	$_{17}$Cl	$_{18}$Ar
$_{19}$K	$_{20}$Ca	$_{31}$Ga	$_{32}$Ge	$_{33}$As	$_{34}$Se	$_{35}$Br	$_{36}$Kr
$_{37}$Rb	$_{38}$Sr	$_{49}$In	$_{50}$Sn	$_{51}$Sb	$_{52}$Te	$_{53}$I	$_{54}$Xe
$_{55}$Cs	$_{56}$Ba	$_{81}$Tl	$_{82}$Pb	$_{83}$Bi	$_{84}$Po		

Notice that within each family of elements the ionization energy decreases as the atomic number increases—that is, it decreases as one proceeds from top to bottom. (This rule is absolutely correct only for the representative elements and some, but not all, transition-element families.)

The values of the ionization energies have been taken relative to that of the ionization energy of hydrogen. Since the main point is to stress regularity and periodicity, it matters very little which units are used, and there is no need to complicate things by using either kcal/mole or electron volts.

It generally takes a long time before students have an adequate "feel" for the concept of energy. It is highly unlikely that they will feel comfortable with discussions of energy at this stage of the game. Nevertheless, unless one is to state authoritatively that "alkali metals lose electrons and halogens gain them," some appeal must be made to the experimental evidence of ionization energies. The emphasis need not be strong, however.

One problem is that if the student has never stopped to think about what is going on, the statement "the ionization energy is high, therefore the tendency to lose electrons is low" or "the ionization energy is low, therefore the tendency to lose electrons is high" may seem to be contradictory. An analogy of any inverse relationship may help here. For example, the *lower* the temperature, the *more* clothes you have to wear to be comfortable.

Postlab Discussion Have one student put his or her arranged data on the chalkboard as a basis for discussion, or use an overhead transparency of this chopped-up, glued-together product. Emphasize the periodicity of properties demonstrated by ionization energy. Point out the arrangement of elements into similar groups vertically and different groups horizontally. Do your students see the trends shown in the chart?

The importance of this miniexperiment and the postlab discussion is to point out the periodicity of the properties of elements and to show the trends. You may wish to examine the periodic trends in other properties, such as atomic and ionic radii, electron affinity, etc. Although polarity of chemical bonds has been discussed in *Reactions and Reason: An Introductory Chemistry Module,* the concept of electronegativity is not mentioned in this module until section *I-24*. Nevertheless, you might wish to discuss it here in connection with ionization energies

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*The *Materials* list for each experiment and miniexperiment in this module is planned for a class of 30 students working in pairs, unless otherwise noted. You may have to adjust this to fit the size of your class.

and electron affinities and show how electronegativity varies periodically.

I-5 CHEMISTRY'S CRYSTAL BALL

The terms *family* and *group* have been used here essentially interchangeably. The older terminology was to speak of Group I, for example, which contained both Family IA, or Subgroup IA (the alkali metals), and Family IB, or Subgroup IB (the coinage metals). It now appears that there is little benefit to be gained by relating the A and B families to each other. About all these two families have in common is a supposed +1 oxidation state, and it is the most stable only for Ag; Cu^{2+} and Au^{3+} are more stable than Cu^{+} and Au^{+}.

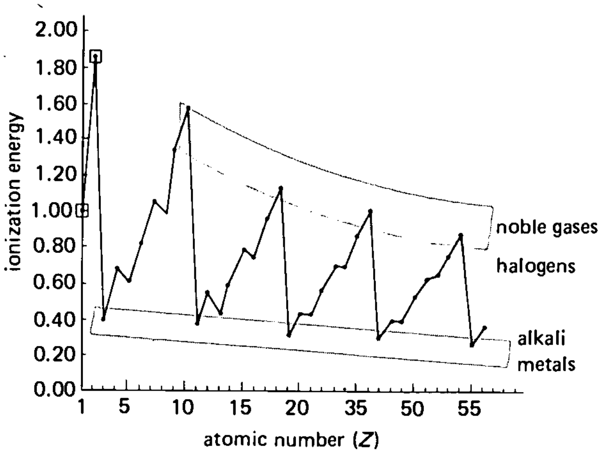
For the same reason, not much is to be gained beyond a little chemical history by discussing Mendeleev's "short form" of the periodic table. Mendeleev used it successfully to predict the properties of both scandium (ekaboron) and gallium (ekaaluminum) and to distinguish these elements (IIIB and IIIA) successfully, but he was intimately familiar with the large differences and what are only formal similarities between the A and B subgroups. For those beginning in chemistry, the short form of the chart is apt to overemphasize nonexistent similarities.

At this point you may wish to touch on the independent discovery of the periodic law by the German chemist Lothar Meyer.* This is an opportunity to remind students that the phenomenon of duplication of ideas and theories occurs frequently in science when enough events point toward a pattern and several independent workers make the same discovery. This process of independent discovery lends credence to the validity of scientific results and ideas. Interestingly enough, although their discoveries were made independently, both Meyer and Mendeleev were associated for a few years with the University of Heidelberg, during a time of remarkable scientific activity there.

I-6 THE HYDROGEN AND HELIUM PUZZLE

Hydrogen and helium do not fit in with the rest of the elements because they cannot form a stable octet of electrons. Helium, however, shares all other properties of the noble gases: high ionization energies, reluctance to form compounds, existence as atoms (rather than diatomic molecules) in gases, etc. The difficulty with placing hydrogen is that it is either an electron donor or acceptor (most often sharing an electron) and may be variously placed according to the whims of the chart-maker.

A possible way to conduct the discussion of this section would be to use an overhead transparency of a periodic chart with the spaces for hydrogen and helium omitted. You can then draw them in with a marker or drop on an overlay with hydrogen and helium on it. Let students suggest various possibilities of placement and try them out on the overhead projector. Encourage the class to respond to them. Point out that because of its ionization energy, helium fits in very nicely with the other elements in Group VIIIA. Hydrogen, on the other hand, exhibits an ionization energy too high to be placed in group IA and too low to be placed in group VIIA.



Although it is common (and perfectly proper) to emphasize the difference between hydrogen and the elements of family IA (lithium, sodium, potassium, etc.) by the complete lack of metallic properties of hydrogen, you should be aware that scientists in this country and the Soviet Union are trying to make "metallic hydrogen" by subjecting the gas to extremely high pressures. Theory predicts that hydrogen should become metallic at fantastically high pressure (millions of atmospheres) [*Science*, 180:398 (27 April 1973)].

I-7 KEEPING UP WITH THE TRENDS

At this point, it is not necessary to delve into

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*See A. J. Ihde, *The Development of Modern Chemistry* (New York: Harper & Row, 1964), pp. 249-251.

the intricacies of the properties of transition metals. The students should be aware that (1) transition metals exist; (2) they have both metallic and nonmetallic properties—that is, Cr and Mn act as metals by forming positive ions, but also as nonmetals by forming anions such as MnO_4^- and $Cr_2O_7^{2-}$; (3) they are placed in the periodic table between representative metals and nonmetals.

A discussion of transition-metal oxidation states is presented later. The only reason transition metals are discussed at all at this point is to show the student that the main trends of the periodic chart remain the same if the transition metals are included.

The intricacies involved in the presence of *d* electrons in the transition metals (and of *f* electrons in the lanthanides and actinides) have been glossed over here, as has the fact that a d^{10} configuration is a pseudo-noble gas configuration. This is in line with our philosophy expressed in both *Reactions and Reason: An Introductory Chemistry Module* and this module that detailed electron configurations (*s, p, d, f,* etc.) do not promote a better understanding of chemistry at the high-school level, especially as usually encountered in the "plug-them-in, crank-them-out" approach to electron configurations. This is especially pronounced in the chemistry of the transition metals, where the stable oxidation states bear little resemblance to the ground state electron configuration. (Have you ever tried to explain to a student why $_{24}Cr$, . . . $3d^5 4s^1$, forms a stable ion, Cr^{3+}, instead of Cr^+? Or why are the 4*s* electrons lost first, even though according to the Aufbau scheme they are not the last ones to be placed?)

Miniexperiment By now, your students have had some exposure to the punning possibilities in chemistry and are acquainted with the names of many of the elements. You may wish to make an overhead transparency of the cartoon map on the following page. Show it to the class and ask the question, "Why are there so many names like europium, americium, francium, and germanium? Surely these are not just bits of word play?"

If a student should ask about the origin of the names of the elements, be prepared to relate some of these names to their origins: *Europium*, discovered by a Frenchman, was named after, obviously, Europe. The elements *francium, germanium,* and *polonium* (Poland) were named by the discoverers after their homelands. A similar name, but less obvious at first glance, is *scandium* (Scandinavia). Even more obscure, unless you are versed in the Latin language and ancient geography, are gallium (L. *Gallia* = France, approximately), rhenium (L. *Rhenus,* the Rhine), hafnium (L. *Hafnia,* Copenhagen), and ruthenium (L. *Ruthenia,* a province of Russia).

A discussion of these names can lead into the broader discussion of the discovery of the elements and the significance of their assigned names. Some elements have been named after cities (holmium, after Stockholm); states (californium); people (mendelevium, curium, fermium, and einsteinium); the colors of their compounds or the element itself: chromium (Gr. *chroma,* color), rhodium (Gr. *rhodon,* rose; cf. rhododendron), and chlorine (Gr. *chloros,* greenish-yellow). For the derivation of the names of the elements, see the *Handbook of Chemistry and Physics*, Chemical Rubber Publishing Company, Cleveland, Ohio. *The Discovery of the Elements*, by Mary Elvira Weeks and Henry M. Leicester, Chemical Education Publishing Company, Easton, Pennsylvania, gives interesting accounts of how the elements were discovered and named.

I-8 FAMILY RELATIONSHIPS

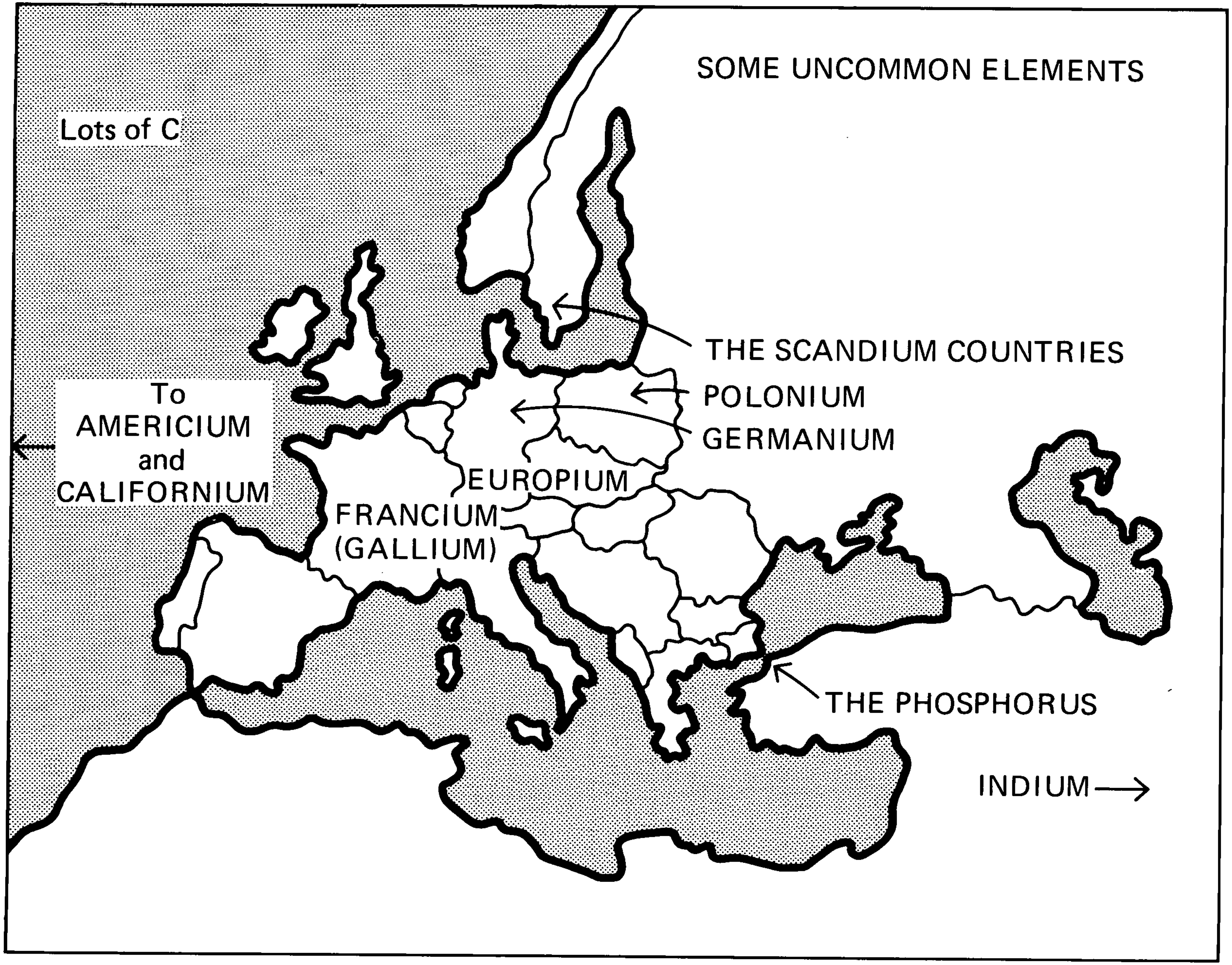
It is difficult to choose good examples for this section. Almost all of the properties of the elements reflect their periodic nature in one way or another, but sometimes the periodicity tends to become obscured. (To extend the previous analogy, there are some times of the year, on vacation trips, for example, when the days of the week tend to become less important.)

There are few good *chemical* examples that may be offered to the student without the addition of many qualifications, explanations, or extensive discussions. For example, in *Chemical Families,* the film mentioned previously, hydrogen is separated from the halogens (X_2) on the basis of two properties:

$$NaH + H_2O \longrightarrow NaOH + H_2$$
$$NaX + H_2O \longrightarrow \text{no reaction (just dissolves)}$$

and

$$P_4 + H_2 \longrightarrow \text{no reaction}$$
$$P_4 + X_2 \longrightarrow P_2X_4,\ PX_3,\ \text{and } PX_5$$


SOME UNCOMMON ELEMENTS
Lots of C
To AMERICIUM and CALIFORNIUM
THE SCANDIUM COUNTRIES
POLONIUM
GERMANIUM
EUROPIUM
FRANCIUM (GALLIUM)
THE PHOSPHORUS
INDIUM

The first distinction is a very good one—the hydride ion is considerably different in its properties from the halide ions. The second distinction is somewhat artificial, however, since, although it correctly distinguishes hydrogen from the halogens, it has been carefully chosen to illustrate the desired difference in properties. First, phosphorus does form compounds with hydrogen, although they cannot be prepared by direct synthesis from the elements. Secondly, if, in place of phosphorus, sulfur had been chosen as a "typical" nonmetal, it would have been found that hydrogen reacts with sulfur (at 600°C), as do fluorine, chlorine, and bromine, but that *iodine does not react*. Using this reaction as a criterion, hydrogen, fluorine, chlorine, and bromine would form a natural family, and iodine would be different!

This is not meant to indicate that there is anything abnormal about sulfur (or iodine) or that the periodic properties of the elements are not overwhelmingly well-documented. It is simply a matter of fact that in most cases the periodicity of chemical properties is based on a large number of reactions and properties, most of which are determined by several factors that tend to add and subtract in their effect. These factors (atomic size, electronegativity, availability of *d* orbitals, etc.) are, in general, too involved to bring into an elementary discussion.

The lanthanide and actinide elements have been given short shrift, beyond an explanation of where these blocks of elements are located. There is sufficient material of more immediate interest that must be included rather than afford the lanthanides and actinides special coverage.

For advanced or able students, it may be desirable to go more deeply into the periodicity of the elements. A nice comparison can be made between the actinides and lanthanides. Until 1944 the elements thorium, protactinium, and uranium were considered to be congeners of hafnium, tantalum, and tungsten. The first attempts to make the element number 93 (later to be named neptunium) were predicated on its supposed similarity to rhenium. The attempts failed. It was about this time that the eminent American chemist Glenn Seaborg suggested that these elements constituted a second inner transition series, analogous to the lanthanides. The Manhattan Project during World War II thus spent a great deal of time and money elucidating the properties of the lanthanides, as well as the synthetically produced actinides. For a more nearly complete discussion of this problem, see G. T. Seaborg, *Man-Made Transuranium Elements* (Englewood Cliffs, N.J.: Prentice-Hall, 1963), chapter 3.

A second topic that might be of interest to students who are already somewhat familiar with the periodic chart is the current interest in synthesizing new elements. The actinide series was completed with element number 103, lawrencium. The two succeeding elements, rutherfordium and hahnium, numbers 104 and 105, are true congeners of hafnium and tantalum. There are exciting possibilities for forming elements in the region of element number 114 and perhaps even higher. See *The Heart of Matter: A Nuclear Chemistry Module* and G. T. Seaborg, "Prospects for Further Considerable Extension of the Periodic Table," *Journal of Chemical Education* (October 1969), pages 626–634.

Note that there is currently some controversy over the discovery of these heavy elements, and both the American group at Berkeley and the Russian group at Dubna have proposed names based on the right of prior discovery. The equivalents are:

Atomic Number	*American Name*	*Russian Name*
102	nobelium	joliotium
103	lawrencium	"103"
104	rutherfordium	kurchatovium
105	hahnium	nielsbohrium
106	*	*

*As yet no name has been proposed.

For your own information concerning some of the "details" of the periodic chart (please don't try to introduce *all* the ideas to your students), you may be interested in reading "Anomalous Properties of Elements That Follow 'Long Periods' of Elements," by J. E. Huheey and C. L. Huheey, *Journal of Chemical Education* (April 1972), pages 227–230.

It is interesting to note that phosphorus was the earliest element to be "discovered"—that is, to be found by experimentation and not just known as common knowledge from the ancients.

It was discovered by Brand in 1669.* In contrast, the first artificially prepared element was technetium, created by the bombardment of molybdenum with neutrons and announced by Perrier and Segre in 1937. The last blank within the periodic chart (this does not include the trans-lawrencium elements, which go on indefinitely) was filled in 1945 by the artificial production of promethium by Marinsky and Glendenin. The *Periodic Table of the Elements* on the following page gives some idea of the dates of discovery of the elements (see also the figure on page 34 of *Reactions and Reason: An Introductory Chemistry Module* for additional information).

MINIEXPERIMENT
I-9 ALL IN THE FAMILY

A demonstration of the periodic trends in chemical reactivity would require a sophisticated setup such as that used in the film *Chemical Families.* Some reactions, however, are rather simple and well within the realm of possibility for most high-school laboratories. This experiment is an example.

Concepts

- There are observable trends in the reactivities of elements in a given family.
- Halogens with greater oxidizing ability will displace other halide ions.

Objective

- Determine relative chemical reactivity of the halogens through redox reactions.

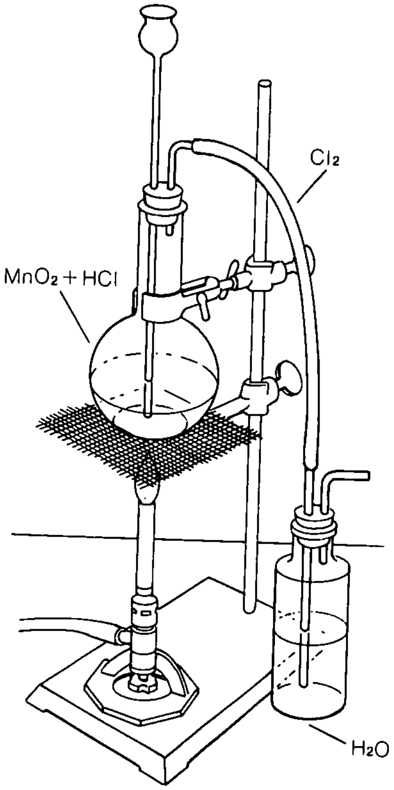
Estimated Time One-half period

Student-grouping Pairs

Materials

200 cm^3 TTE, trichlorotrifluoroethane (Freon 113 or Du Pont TF Solvent)
50 cm^3 chlorine water
50 cm^3 bromine water
50 cm^3 iodine solution
50 cm^3 0.1 M KCl
50 cm^3 0.1 M KBr
50 cm^3 0.1 M KI
90 18 × 150-mm test tubes
90 test-tube corks
15 test-tube racks
15 10-cm^3 graduated cylinders

Advance Preparation *Chlorine Water*—For convenience, commercial bleach, such as Clorox and Purex, acidified with concentrated HCl (4–6 drops concentrated HCl per 100 cm^3 bleach) can be labeled as "chlorine water" and used in the experiment. Chlorine water may be used, if you have it, or it may be prepared by bubbling chlorine gas into a flask containing distilled water. **Caution:** *Cl_2 is poisonous. This operation should be conducted under a hood in a well-ventilated room.* If a tank of Cl_2 is not available, chlorine may be generated by the oxidation of HCl (see illustration); 10 g MnO_2 and 40 cm^3 concentrated HCl will generate sufficient Cl_2. Heat the reaction vessel gently. Cl_2 will be collected in the cold water bottle.



Optional setup for generation of chlorine gas

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*See M. E. Weeks and H. M. Leicester, *Discovery of the Elements,* 7th ed. (Easton, Pa: Chemical Education Publishing Company, 1968).

PERIODIC TABLE OF THE ELEMENTS

1.008
H
Hydrogen
1

200.6 — atomic mass
Hg — symbol
Mercury — name
80 — atomic number

IA	IIA	IIIB	IVB	VB	VIB	VIIB	← VIIIB →			IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
																	4.00 **He** Helium 2
6.94 **Li** Lithium 3	9.01 **Be** Beryllium 4											10.8 **B** Boron 5	12.0 **C** Carbon 6	14.0 **N** Nitrogen 7	16.0 **O** Oxygen 8	19.0 **F** Fluorine 9	20.2 **Ne** Neon 10
23.0 **Na** Sodium 11	24.3 **Mg** Magnesium 12											27.0 **Al** Aluminum 13	28.1 **Si** Silicon 14	31.0 **P** Phosphorus 15	32.1 **S** Sulfur 16	35.5 **Cl** Chlorine 17	39.9 **Ar** Argon 18
39.1 **K** Potassium 19	40.1 **Ca** Calcium 20	45.0 **Sc** Scandium 21	47.9 **Ti** Titanium 22	50.9 **V** Vanadium 23	52.0 **Cr** Chromium 24	54.9 **Mn** Manganese 25	55.8 **Fe** Iron 26	58.9 **Co** Cobalt 27	58.7 **Ni** Nickel 28	63.5 **Cu** Copper 29	65.4 **Zn** Zinc 30	69 7 **Ga** Gallium 31	72.6 **Ge** Germanium 32	74.9 **As** Arsenic 33	79.0 **Se** Selenium 34	79.9 **Br** Bromine 35	83 8 **Kr** Krypton 36
85.5 **Rb** Rubidium 37	87.6 **Sr** Strontium 38	88.9 **Y** Yttrium 39	91.2 **Zr** Zirconium 40	92.9 **Nb** Niobium 41	95.9 **Mo** Molybd'm 42	(97) **Tc** Technetium 43	101.1 **Ru** Ruthenium 44	102.9 **Rh** Rhodium 45	106.4 **Pd** Palladium 46	107.9 **Ag** Silver 47	112.4 **Cd** Cadmium 48	114.8 **In** Indium 49	118.7 **Sn** Tin 50	121.8 **Sb** Antimony 51	127.6 **Te** Tellurium 52	126.9 **I** Iodine 53	131 3 **Xe** Xenon 54
132.9 **Cs** Cesium 55	137.3 **Ba** Barium 56	138.9 **La** * Lanthanum 57	178.5 **Hf** Hafnium 72	180.9 **Ta** Tantalum 73	183.8 **W** Tungsten 74	186.2 **Re** Rhenium 75	190.2 **Os** Osmium 76	192.2 **Ir** Iridium 77	195.1 **Pt** Platinum 78	197.0 **Au** Gold 79	200.6 **Hg** Mercury 80	204.4 **Tl** Thallium 81	207.2 **Pb** Lead 82	209.0 **Bi** Bismuth 83	(209) **Po** Polonium 84	(210) **At** Astatine 85	(222) **Rn** Radon 86
(223) **Fr** † Francium 87	226 0 **Ra** Radium 88	(227) **Ac** ** Actinium 89	‡ 104	‡ 105	‡ 106												

*	140.1 **Ce** Cerium 58	140 9 **Pr** Praseody'm 59	144 2 **Nd** Neodymium 60	(145) **Pm** Promethium 61	150 4 **Sm** Samarium 62	152.0 **Eu** Europium 63	157 3 **Gd** Gadolinium 64	158.9 **Tb** Terbium 65	162 5 **Dy** Dysprosium 66	164 9 **Ho** Holmium 67	167.3 **Er** Erbium 68	168 9 **Tm** Thulium 69	173 0 **Yb** Ytterbium 70	175.0 **Lu** Lutetium 71
**	232.0 **Th** Thorium 90	231 0 **Pa** Protactinium 91	238.0 **U** Uranium 92	237.0 **Np** Neptunium 93	(242) **Pu** Plutonium 94	(243) **Am** Americium 95	(245) **Cm** Curium 96	(245) **Bk** Berkelium 97	(251) **Cf** Californium 98	(254) **Es** Einsteinium 99	(254) **Fm** Fermium 100	(256) **Md** Mendel'm 101	(254) **No** Nobelium 102	(257) **Lr** Lawrencium 103

†The most stable known isotopes are shown in parentheses.

‡The discovery of elements 104, 105, and 106 has been claimed by both American and Soviet scientists. The Americans have suggested the name *rutherfordium* and *hahnium* for 104 and 105; the Soviets have suggested the names *kurchatovium* and *nielsbohrium* for these same elements. No name has yet been proposed for element 106.

Elements discovered ☐ or artificially made since 1870.

27

Bromine Water—If bromine water as such is not available, dissolve 4 grams KBr in 50 cm^3 chlorine water or commercial bleach (Clorox or Purex) and label this as "bromine water." If elemental bromine is available, pour the bromine into a flask containing distilled water, stir, let settle, and decant bromine water. The excess Br_2 will simply settle to the bottom. Fresh bromine water can then be prepared by adding more water to the flask.

Iodine Solution—The iodine solution (I, in KI reagent) may be prepared by dissolving 1 g KI in distilled water, adding 0.5 g I_2, and diluting to 50 cm^3.

To prepare 0.1 M KCl, dissolve 0.4 g in 50 cm^3 water.

To prepare 0.1 M KBr, dissolve 0.6 g KBr in 50 cm^3 water.

To prepare 0.1 M KI, dissolve 0.8 g KI in 50 cm^3 water.

Prelab Discussion See section *I-8*.

Laboratory Tips To facilitate handling of the chemicals, place solutions in small bottles and use medicine droppers to disperse. **Caution:** *The halogen solutions are corrosive and should be flushed away with water immediately if they come in contact with the skin.*

Postlab Discussion On the basis of their observations, students should complete and balance the equations in the test:

$$2KBr + Cl_2 \longrightarrow 2KCl + Br_2$$
$$2KI + Cl_2 \longrightarrow 2KCl + I_2$$
$$KCl + Br_2 \longrightarrow \text{N.R.}$$
$$2KI + Br_2 \longrightarrow 2KBr + I_2$$
$$KCl + I_2 \longrightarrow \text{N.R.}$$
$$KBr + I_2 \longrightarrow \text{N.R.}$$

In discussing the "missing element," fluorine, let your students come to the realization that since F_2 would be at the top, it would react with all the other halides. In addition, manufacturing F_2 requires special methods and it is so reactive that it is a dangerous chemical to have in the laboratory.

The students may also inquire into the absence of astatine. This element is very rare and radioactive. Studies indicate that it behaves as expected and is less reactive than iodine.

I-10 GAINERS AND LOSERS

The point to stress here is the stable octet of electrons possessed by the Group VIIIA elements. The halogens, Group VIIA, have one electron missing from their octet and therefore tend to accept electrons:

$$X + e^- \longrightarrow X^-$$

The alkali metals have an electron beyond the stable octet and tend to lose that electron:

$$M \longrightarrow M^+ + e^-$$

In contrast to section *I-8*, where we considered trends *within* families, here the emphasis is on differences *between* families.

The students may benefit by using a worksheet for writing formulas of binary compounds. Keep the majority of examples brief and similar to the examples in the module.

Your students may benefit from a discussion of the periodic table as arranged by Glenn T. Seaborg. If you wish, prepare a transparency of the periodic table illustrated on the following page for use on the overhead projector. This duplicates the smaller version in the student module.

You may wish to use this table in conjunction with the Mendeleev tables and the present-day periodic table as a review of this section.

ANSWERS TO PROBLEMS

(Student module page 24)

1. (Cl, Br, I); (Ca, Sr, Ra); (Cu, Ag, Au)
2. Cl, Br, and I have similar properties and form similar compounds, as seen in *I-9* and *I-10*. Cu, Ag, and Au are all metals and have all been used to make coins; in fact, the group is known as the "coinage metals." Ca is a major component of the structure of bone. Both radioactive strontium (^{90}Sr) and radium (^{226}Ra) can cause bone disease, the former from radioactive fallout and the latter from a form of radium poisoning exemplified in the notorious case of women who worked in factories painting luminous watch dials.
3. Al, Ga, In
4. Cs, Mo, Ca, K, Fe, and Ni are metals; C, S, N, I, B, P, and Ne are nonmetals.
5. $2KBr + Cl_2 \longrightarrow 2KCl + Br_2$
   $2KI + Br_2 \longrightarrow 2KBr + I_2$
   $2KCl + F_2 \longrightarrow 2KF + Cl_2$

PERIODIC TABLE SHOWING HEAVY ELEMENTS AS MEMBERS OF AN ACTINIDE SERIES

ARRANGEMENT BY GLENN T. SEABORG

1945

1 H 1.008																1 H 1.008	2 He 4.003
3 Li 6.940	4 Be 9.02											5 B 10.82	6 C 12.010	7 N 14.008	8 O 16.000	9 F 19.00	10 Ne 20.183
11 Na 22.997	12 Mg 24.32	13 Al 26.97										13 Al 26.97	14 Si 28.06	15 P 30.98	16 S 32.06	17 Cl 35.457	18 A 39.944
19 K 39.096	20 Ca 40.08	21 Sc 45.10	22 Ti 47.90	23 V 50.95	24 Cr 52.01	25 Mn 54.93	26 Fe 55.85	27 Co 58.94	28 Ni 58.69	29 Cu 63.57	30 Zn 65.38	31 Ga 69.72	32 Ge 72.60	33 As 74.91	34 Se 78.96	35 Br 79.916	36 Kr 83.7
37 Rb 85.48	38 Sr 87.63	39 Y 88.92	40 Zr 91.22	41 Cb 92.91	42 Mo 95.95	43	44 Ru 101.7	45 Rh 102.91	46 Pd 106.7	47 Ag 107.880	48 Cd 112.41	49 In 114.76	50 Sn 118.70	51 Sb 121.76	52 Te 127.61	53 I 126.92	54 Xe 131.3
55 Cs 132.91	56 Ba 137.36	57 La 138.92 58-71 See La series	72 Hf 178.6	73 Ta 180.88	74 W 183.92	75 Re 186.31	76 Os 190.2	77 Ir 193.1	78 Pt 195.23	79 Au 197.2	80 Hg 200.61	81 Tl 204.39	82 Pb 207.21	83 Bi 209.00	84 Po	85	86 Rn 222
87	88 Ra	89 Ac See Ac series	90 Th	91 Pa	92 U	93 Np	94 Pu	95	96								

LANTHANIDE SERIES	57 La 138.92	58 Ce 140.13	59 Pr 140.92	60 Nd 144.27	61	62 Sm 150.43	63 Eu 152.0	64 Gd 156.9	65 Tb 159.2	66 Dy 162.46	67 Ho 163.5	68 Er 167.2	69 Tm 169.4	70 Yb 173.04	71 Lu 174.99
ACTINIDE SERIES	89 Ac	90 Th 232.12	91 Pa 231	92 U 238.07	93 Np 237	94 Pu	95	96							

EVALUATION ITEMS

These are evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. Metallic elements are normally

   A. good conductors of electricity and malleable.
   B. good conductors of electricity and brittle.
   C. poor conductors of electricity and malleable.
   D. poor conductors of electricity and brittle.

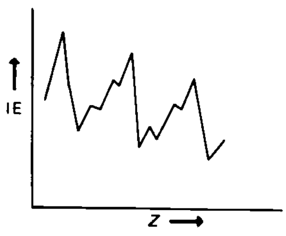
2. As one moves along a given row in the periodic table, ionization energy

   A. decreases from left to right.
   B. increases from left to right.
   C. remains unchanged.
   D. increases, then decreases.

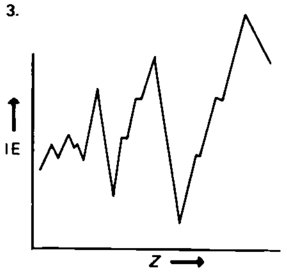
3. Which of the following diagrams best represents the relationship between ionization energy and atomic number?

   A. 1 B. 2 C. 3 D. 4

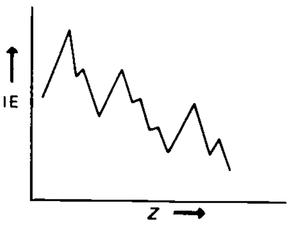
1.



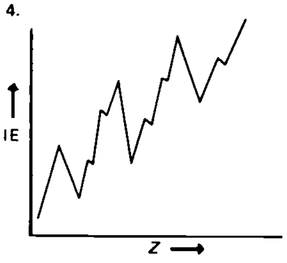
3.



2.



4.



4. Which of the following explanations is best for not placing hydrogen in either group IA (alkali metals) or group VIIA (halogens)?

   A. Hydrogen is a much lighter element than the alkali metals or the halogens.
   B. Hydrogen can form compounds with all other elements.
   C. The ionization energy of hydrogen is too high for group IA, but too low for group VIIA.
   D. None of the above

5. Which of the following statements is true?

   A. Transition metals are poor conductors of electricity.
   B. Transition metals are found between the Group IIA metals and the Group IIIA metals and nonmetals.
   C. Transition metals do not exhibit periodicity in their chemical and physical properties.
   D. Transition metals do not exhibit both metallic and nonmetallic properties.

6. The most chemically active of the elements I, Cu, He, and F is:

   A. I C. He
   B. Cu D. F

7. The least chemically active of the elements F, Cl, Br, and I is:

   A. F C. Br
   B. Cl D. I

8. Which element has the greatest tendency to lose electrons?

   A. Fr C. S
   B. F D. Be

9. Which of the following would be a correct formula for a compound between H and F?

   A. H_2F C. H_2F_3
   B. HF_2 D. None of the above

Structural Chemistry of Metals and Their Compounds

By now students are aware of the chemist's reliance upon an understanding of periodic properties. Here the properties of metals are explained in relation to their lattice structures. Why metals useful in construction are largely members of the transition elements is explained by analogy, using models of their lattice structures. Coordination numbers, packing efficiencies, and types of lattice structures are discussed and studied through experiments.

A similar study is made of some representative ionic compounds. Opportunities to contrast the properties of metallic and ionic solids are provided by considering the nature of their lattice structures as well as electrical charge or absence of charge. (For instance, why are metals malleable and ionic crystals brittle?)

I-11 WHY ARE METALS "METALLIC"?

In some ways, metals are very simple materials. Elemental metals such as gold are composed of only one type of atom, and their geometry (as becomes apparent in the following sections) is restricted to a few simple types. On the other hand, the bonding in metals is difficult to describe in simple, nonquantum-mechanical terms. In this section, a short introduction to the problem is stated with a provisional explanation of the bonding: The metallic bond resembles Ping-Pong balls held together by molasses—that is, strong attractive forces ("stickiness"), but there is freedom for the atoms to move past each other (metals are malleable), and the electrons ("molasses") can flow as well (electric current). These properties are discussed in more detail later, so do not feel compelled to go into the subject at great length now. The section is meant to be exactly what its title implies—a question. As the students perform the experiments and miniexperiments, facts will be learned. They will develop theories that will reinforce an understanding of metals.

MINIEXPERIMENT
I-12 WHO'S WHO?

The purpose of this miniexperiment is to acquaint the student with some of the important properties of metals and nonmetals and to use these properties to classify elements into the two groups.

Concepts

- Metals have characteristic and distinctive properties.
- Nonmetals have characteristic and distinctive properties.

Objective

- Classify elements as metals and nonmetals on the basis of their physical properties.

Estimated Time One-half period

Student Grouping Pairs

Materials

15 10-cm lengths of:
  copper wire
  aluminum wire
  lead wire
  zinc wire
  tungsten wire
  molybdenum wire

Note: Many metals in the form of wire (lead and tungsten, for example) may be purchased from Research Organic/Inorganic Chemical Corporation, 11686 Sheldon Street, Sun Valley, California 91352, or from Alfa Division, Ventron Corporation, P.O. Box 299, Danvers, Massachusetts 01923.

You will also need:
10 g sodium
5 g lithium
5 g graphite (pencil lead)
5 g iodine
50 g sulfur (lump)
conductivity apparatus (1.5 volt D cell, 1.5 volt bulb, alligator clips, wire, miniature socket)

Best results will be obtained by using wires of comparable diameter. If a stock of wires is not available, use pieces of metal of similar dimension. Copper and zinc sheets are often in laboratory stock. Brass shim stock may be obtained from the school shop, lead from plumber's wool, and tungsten from light-bulb filaments.

There are only a few nonmetals that are solids which can be used for comparison. Lump sulfur (not flowers of sulfur) is good, and iodine can be used if students are warned that it is corrosive. Graphite from lead for mechanical pencils is suitable, although it is an exceptional nonmetal that *does* conduct electricity. White phosphorus could be used in a teacher demonstration, but it is too flammable and poisonous for student use.

Advance Preparation See *Materials*

Prelab Discussion We have purposely left the instructions to the students somewhat nonspecific to allow the teacher some flexibility in choosing materials and tests to be performed. Call to your students' attention the materials that you have selected for testing. Suggest to your students that they test the materials for (1) physical appearance, (2) electrical conductivity, and (3) malleability and ductility.

Students can test for *physical appearance* by observing properties such as color, luster, texture, and shape. To test conductivity, you can set up a simple conductivity apparatus. All you will need is a dry cell, a miniature socket, a flashlight bulb, bell wire, and alligator clips. You can test for malleability and ductility by drawing wires, bending metals, and hammering metals.

Laboratory Tips You might get the class started by conducting some demonstrations. For example, you can easily demonstrate the low melting point (70°C) of Wood's Metal (50 percent bismuth, 25 percent lead, 12.5 percent tin, 12.5 percent cadmium). Teaspoons molded from Wood's Metal melt when used for stirring hot tea. You can demonstrate the physical properties of sodium (luster, malleability, conductivity, relative softness) by cutting a chunk with a knife to expose a fresh surface. Note that the luster quickly fades when a film of oxide forms as a result of contact with the air.

Range of Results The greater the number of samples your class works with, the greater is the number of properties the students will observe. Suggest to your students that they each compile a list of the properties that their tests reveal. Have students compare their lists. In all probability, the lists will vary. If conflicting conclusions arise, have the students retest the material that is in question.

Postlab Discussion Have your students go over the lists of properties that they have compiled. Then compile a general list from the students' individual lists. Put this list on the chalkboard. Call upon a student to enumerate some differences between metals and nonmetals. Invite others in the class to comment on the student's enumeration and, if necessary, to add to his or her list. Which materials are metals? Which materials are nonmetals? Discuss. Form conclusions.

I-13 HOW'S YOUR COORDINATION?

A brief introduction to the concepts of coordination number and efficiency of packing can serve as a prelab discussion to the following experiment, *I-14*. However, it may prove more interesting to plunge right into the experiment and discuss both concepts as the need arises.

EXPERIMENT
I-14 BECOMING AN EFFICIENCY EXPERT

The purpose of this experiment is to acquaint the student with the relationship between coordination numbers of different lattices and their efficiency of packing.

There are no "right" or "wrong" ways to arrange the balls, only different ways, some more efficient than others. The plastic boxes define the boundaries within which the student will "pack" atoms. Plastic boxes as described are recommended for use in this experiment, since substitutes do not have appropriate dimensions.

Although we feel strongly that there is no substitute for "hands on" experience by the students—building models, trying various possibilities, and getting results that can be discussed and even debated—certainly it is also very useful to have prebuilt models available for use. The advantages of prebuilt models include accuracy, stability (nothing is as exasperating as having a structure "collapse" at a critical point in a discussion), clarity, and versatility. They need not be expensive models; models quite suitable for use in your classes can readily be constructed from Styrofoam balls similar to those used in this experiment. Construction of demonstration models can be an interesting special project for some of your students.

Concepts

- The coordination number of an atom in a lattice is determined by counting the number of nearest neighbors.

- Different lattice structures may have different coordination numbers.
- Some lattice structures have greater packing efficiency.

Objectives

- Determine the coordination number of lattice structures.
- Construct three-dimensional metallic lattices, using plastic balls and boxes as models.
- Calculate the packing efficiency of a lattice structure.

Estimated Time One and one-half periods

Student Grouping Pairs

Materials

15 plastic boxes
15 false bottoms
15 sets of Styrofoam or polystyrene balls, 1″ (2.54 cm) diameter (150 to 200 balls in a set)
84 polystyrene balls, 5 cm to 8 cm in diameter (optional)

Enough copies of a worksheet for this experiment may be duplicated ahead of time (see following page).

Note: You can make the plastic boxes out of ⅛″ Plexiglas sheets,* 4′ × 6′ in size. Such sheets can be purchased from any Plexiglas dealer. Styrofoam and polystyrene balls are available from Edmund Scientific Company, 801 Edscorp Building, Barrington, New Jersey 08007, or from Science Related Materials, P.O. Box 1442, Janesville, Wisconsin 53345. Science Related Materials also offers cork balls and precolored polystyrene balls, but they are more expensive.

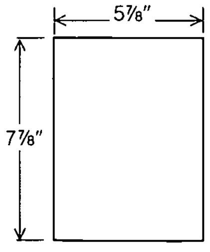
Advance Preparation The plastic boxes are constructed of clear acrylic material (Plexiglas), 6″ × 8″ (inside dimensions of base) and about 3″ high. Use methylene chloride solvent to glue the pieces of plastic together. The false bottoms are single pieces, 5⅞″ × 7⅞″. For the body-centered cubic (BCC) packing activity, you can use strips that measure 5⅞″ × 2½″ (approximately). Instructions for the use of these strips are discussed in the *Laboratory Section*. These dimensions are for the ⅛″ thick Plexiglas. If other thicknesses are employed, the dimensions of the end pieces and side pieces should be adjusted accordingly. The plastic sheet can be cut by using a Plexiglas cutter or saw.

Wet the edges of the pieces with methylene chloride. Apply the methylene chloride at the joint with a syringe or a medicine dropper, with the tip drawn to form a capillary tip. Methylene chloride dries rather quickly, and it therefore will not be necessary to hold the pieces for too long. No clamps will be needed. After gluing the side pieces, add the end pieces in the same manner. Glue the end pieces to the edges of the bottom piece and the side pieces.

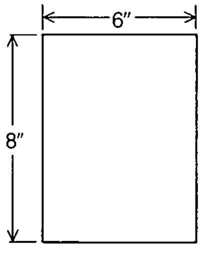
Give each pair or group of students a set of plastic balls of Styrofoam or expanded polystyrene, 1″ in diameter and all of one color (150 to 200 balls in a set). Half of the student groups can work with white balls and half with colored balls. You can color the balls by dipping them in latex paint. Spray paints can also be used, but difficulties sometimes arise as a result of the softening effect of spray paints. We prefer the harder and slicker expanded polystyrene balls over the rough-surfaced Styrofoam.

You will need the following transparencies:

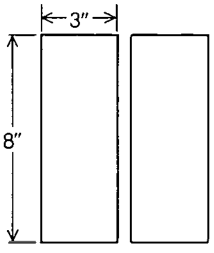
Square lattice (transparency 1)
Hexagonal lattice (transparency 2)



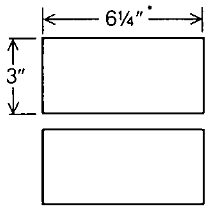
false bottom



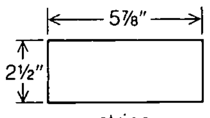
bottom piece



side pieces



end pieces



strips

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*Although we are firmly committed to the metrication of America, we still have a long way to go. Any teacher who goes into the school shop and asks for a 102 × 152-mm box made of 3.175-mm thick Plexiglas is probably asking for trouble. Therefore all dimensions are given in inches.

Worksheet—Experiment *I-14*

1. Calculation of the volume of a single atom:

   A. ___________ (diameter of atom) × ½ = ___________ (radius)

   B. ___________ (radius)^3 ___________ (r^3)

   C. ___________ (r^3) × 1.333 × 3.1416 (pi) = ___________ (volume of each atom)

2. The simple cubic lattice structure:

   A. _______ (no. of atoms) × _______ (vol. of each atom) = _______

   B. _______ (W) × _______ (L) × _______ (H) = _______ (vol. of box)

   $$\text{Efficiency} = \frac{\text{Volume of atoms (2A)}}{\text{Volume of lattice (2B)}} = \_\_\_\_\_\_\_$$

3. The face-centered cubic lattice (FCC) structure:

   A. _______ (no. of atoms) × _______ (vol. of each atom) = _______

   B. _______ (W) × _______ (L) × _______ (H) = _______ (vol. of box)

   $$\text{Efficiency} = \frac{(3A)}{(3B)} = \_\_\_\_\_\_\_$$

   The coordination number of this structure is _______.

4. The body-centered cubic lattice (BCC) structure: (optional)

   A. _______ (no. of atoms) × _______ (vol. of each atom) = _______

   B. _______ (W) × _______ (L) × _______ (H) = _______ (vol. of box)

   $$\text{Efficiency} = \frac{(4A)}{(4B)} = \_\_\_\_\_\_\_$$

   The coordination number of this structure is _______.

5. The hexagonal, closest-packed lattice (HCP) structure:

   A. _______ (no. of atoms) × _______ (vol. of each atom) = _______

   B. _______ (W) × _______ (L) × _______ (H) = _______ (vol. of box)

   $$\text{Efficiency} = \frac{(5A)}{(5B)} = \_\_\_\_\_\_\_$$

   The coordination number of this structure is _______.

6. The cubic closest-packed (CCP) structure:

   A. _______ (no. of atoms) × _______ (vol. of each atom) = _______

   B. _______ (W) × _______ (L) × _______ (H) = _______ (vol. of box)

   $$\text{Efficiency} = \frac{(6A)}{(6B)} = \_\_\_\_\_\_\_$$

   The coordination number of this structure is ___________.

Second layer of hexagonal closest-packing and cubic closest-packing (transparency 3, preferably of another color)
Third layer of hexagonal closest-packing (transparency 2; second copy), where circles are marked with felt-tip pen
Third layer of cubic closest-packing (transparency 4), colored differently
Bottom layer of body-centered cubic (transparency 5)
Second layer of body-centered cubic (transparency 6)

Models for these transparencies can be found on the following pages. Given the extra work involved in making transparencies 5 and 6, the models provided are drawn to actual size. Each transparency should be a different color, either made from colored plastic or marked with a felt-tip pen.

Prelab Discussion Explain the general purpose of the experiment. (The purpose is to establish the different ways that atoms can be arranged in a crystal structure.) Try to avoid giving the impression that there is a "right" way to do it. Encourage individuality. The experiment is most successful if half the class gets one arrangement and half gets another.

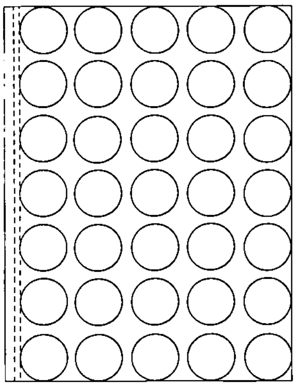
Laboratory Tips A suggested way of conducting this experiment would be to have an overhead projector set up with all the transparencies of the different arrays already prepared and ready to be used.

Have the students measure the dimensions and calculate the efficiency of each lattice as it is built. As a particular structure is completed, you might project the structure on the overhead projector as a transparency, by placing a transparent box directly on the overhead projector, or by using transparencies 1 and 2. Briefly discuss the structures and their differences as they are completed. After different structures have been tried, ask all of the students to make a square array as the bottom layer in their boxes. Then ask them to add two more layers in any way they choose. If the second layer is identical and congruent with the first, your students have constructed a *simple cubic lattice* with C.N. of 6. If the second layer is a square lattice, but displaced so that the atoms lie in the "pocket" between four atoms of the first layer, they have constructed a *face-centered cubic lattice* with C.N. of 12.

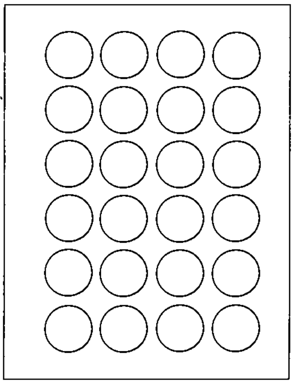
Here, as well as later in the module (page 37), the students are working with two-dimensional arrays before proceeding to three-dimensional lattices. If any of your students object that the "real world" is three-dimensional and that working in two dimensions is a waste of time, you might refer them to an article that discusses monolayers of adsorbed materials as two-dimensional solids, liquids, and gases! See J. C. Dash, "Two-Dimensional Matter," *Scientific American* (April 1973), pages 30–40.

Body-Centered Cubic (Optional). This lattice is almost impossible to make without a fixed starting layer, since the atoms are not tangent in the square layers in the box. Place a copy of the array in transparency 5 in the bottom of a plastic box and place a false bottom over it. Now glue 35 balls over the marked spots. A BCC array does not fit exactly into a 6″ × 8″ box. To increase the accuracy of the model, a 2½″ × 5⅞″ strip can be glued to the false bottom over the dotted lines. The balls should be tangent to the box on all sides now. Have the students make a BCC lattice or, (if only one BCC array is available) make it for them. To project an image of the body-centered cubic lattice, place transparency 5 on the projector stage, and then drop transparency 6 over 5 as an overlay.

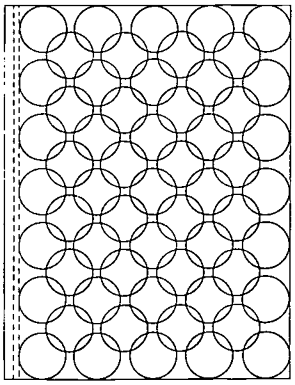
TRANSPARENCY 5



TRANSPARENCY 6

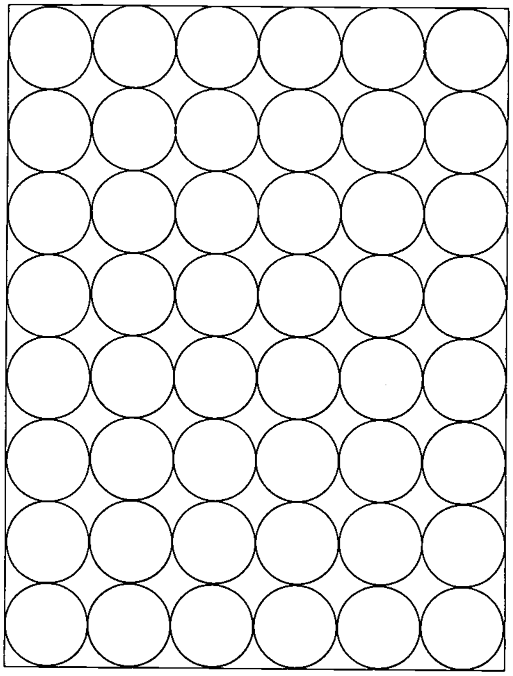


5 WITH 6 AS OVERLAY



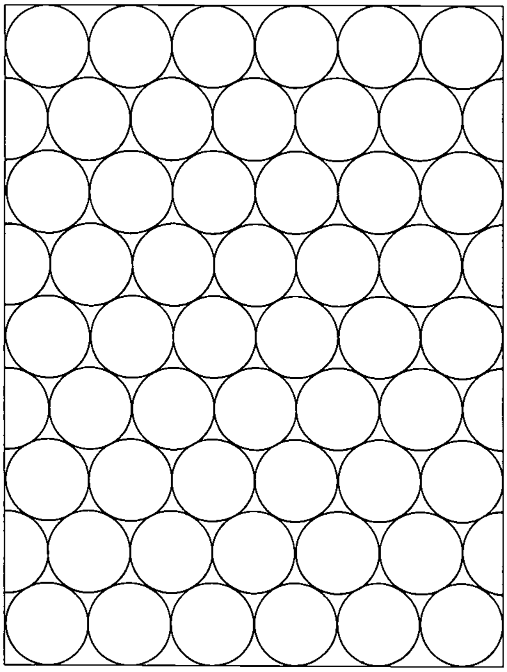
TRANSPARENCY 1

square lattice



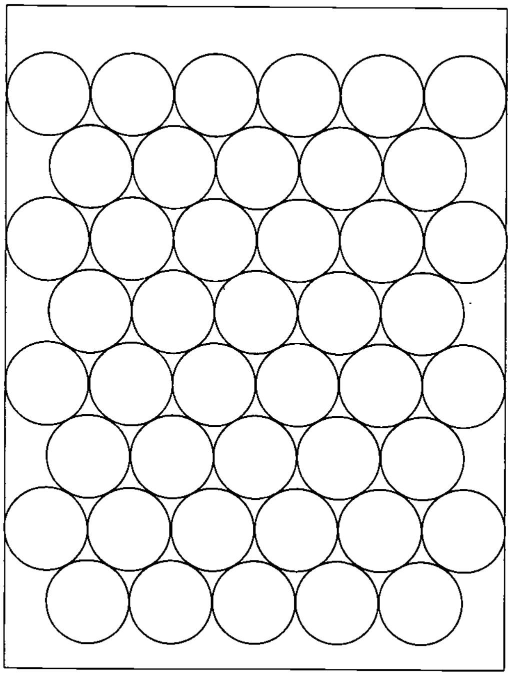
TRANSPARENCY 2

hexagonal lattice



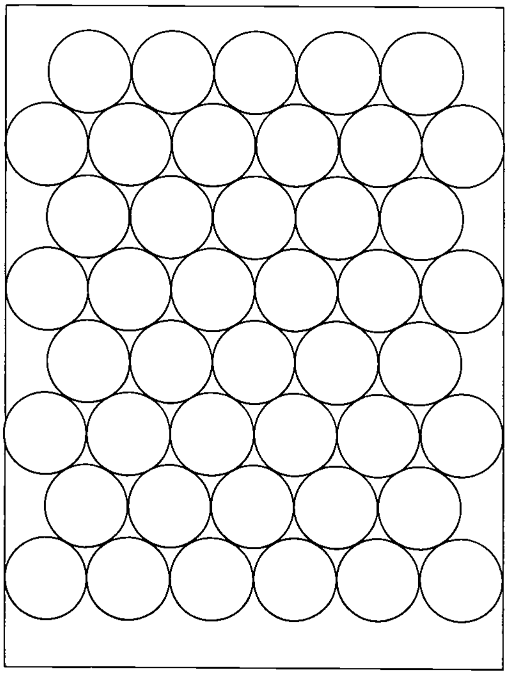
TRANSPARENCY 3

2nd layer of HCP and CCP



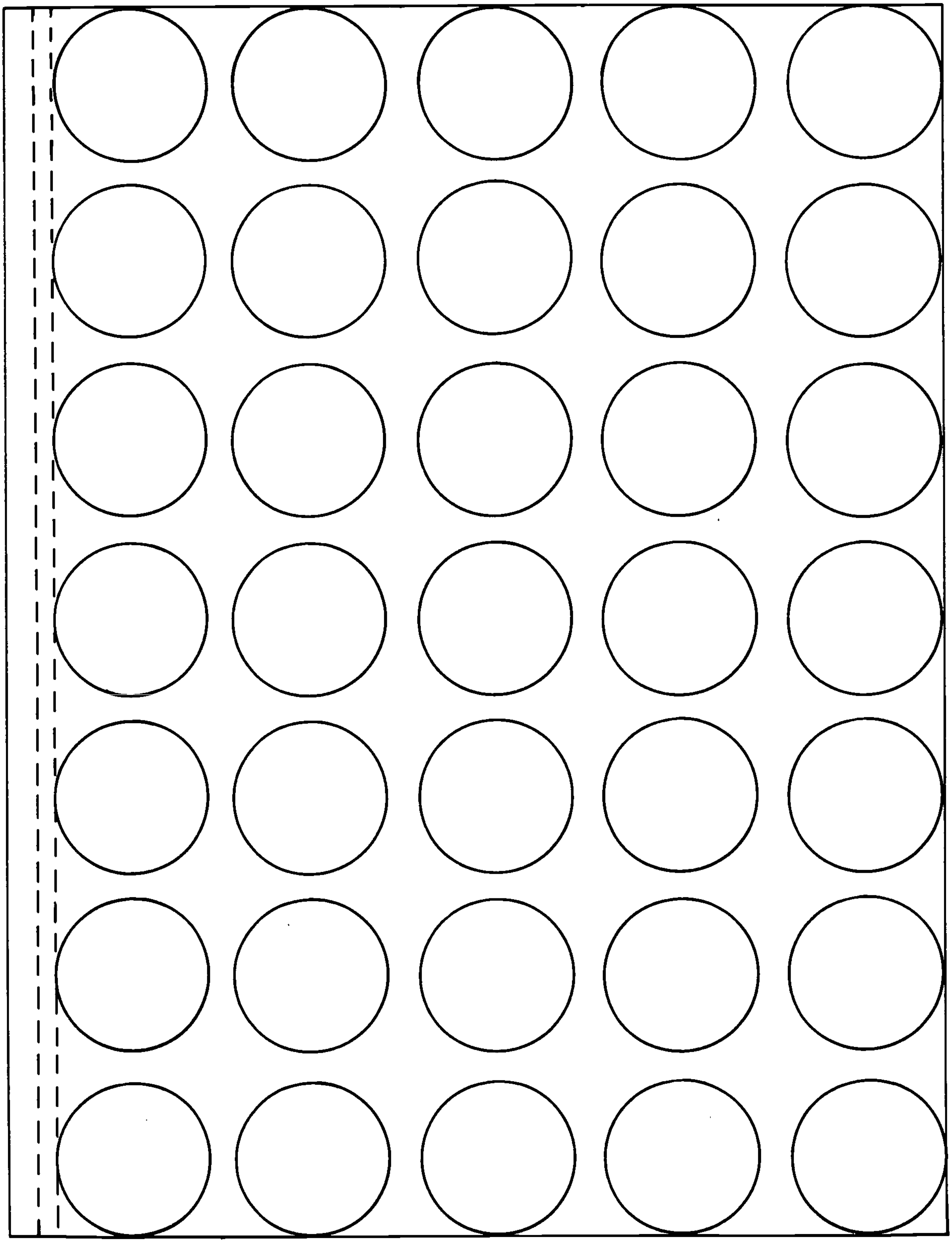
TRANSPARENCY 4

3rd layer of CCP

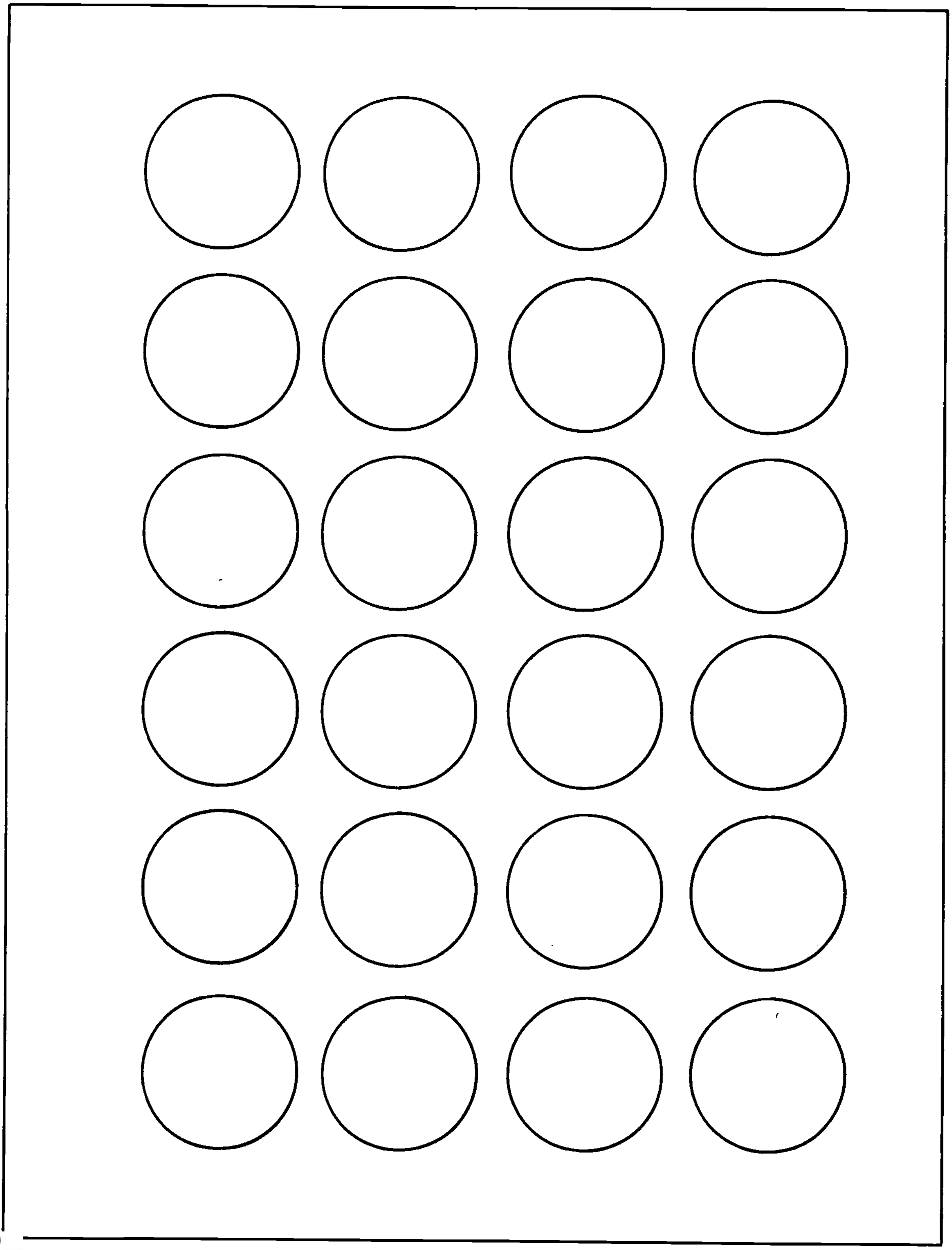


TRANSPARENCY 5

bottom layer of BCC



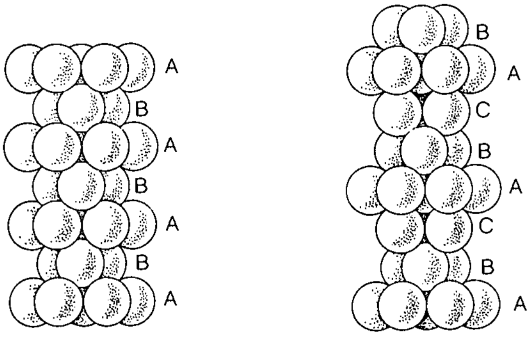
TRANSPARENCY 6
2nd layer of BCC



Next ask the students to make a hexagonal array in the bottom layer of their boxes. Ask them to add two layers in any way they choose.

A. Hexagonal closest-packed lattice, second layer in "pockets" between three atoms of first layer, third layer directly over the atoms in the first, (C.N. = 12). Called an AB, AB, AB structure. You may show a second copy of transparency 2 on top of transparencies 3 and 2. Notice that there are still "holes" in the array.

B. Cubic closest-packed lattice, second layer as in HCP, third not directly above first layer (there is only one way to do this). (C.N. = 12). Called an ABC, ABC structure. Show transparency 4 on top of transparencies 3 and 2.

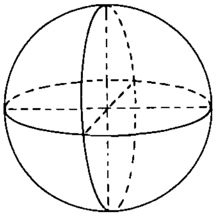


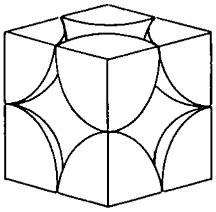
(1) (2)

Arrangement of layers (1) AB, AB, AB in hexagonal closest-packed structure and (2) ABC, ABC in cubic closest-packed structure.

A note in the *Journal of Chemical Education* 49 (1972), page 674, suggests a method of using the overhead projector to help the students see the difference between HCP and CCP.

The determination of the coordination number should offer no special problem. In calculating efficiencies, however, point out to the student the need to isolate, at least mentally, a "unit" cell, in which case one must consider the volume occupied by fractions of balls. To help the student with this you may build a "unit cell" for a simple structure like the simple cubic. Obtain a large Styrofoam ball and slice it in such a way that you get eight octants. Then turn the octants around and connect them with toothpicks in such a way that a cube is formed.





This makes evident the fact that only eight ⅛ balls or one full ball occupy any space in the unit cell of side 2r.

Hence the efficiency in this case is

$$\text{Efficiency} = \frac{(8)(1/8)(4/3\,\pi r^3)}{(2r)^3} \times 100\% = 52.3\%$$

Range of Results The expected values for the different structures are

Lattice Type	*Coordination Number*	*Efficiency*
simple cubic	6	52%
face-centered cubic	12	74%
body-centered cubic	8	68%
hexagonal closest-packed	12	74%
cubic closest-packed	12	74%

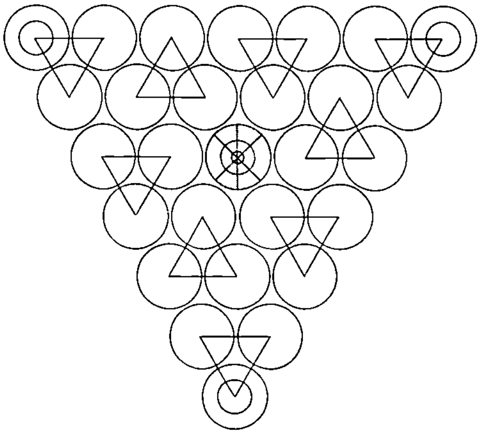
Although most of the students should be able to obtain the correct coordination numbers, the calculated efficiencies (with the exception of the simple cubic) will generally be lower. A calculated value higher than those listed above will probably be the result of a gross error in counting the balls or in the calculations. The low values result from not including the fractions of balls that would fit near the edges of some of the layers.

Postlab Discussion Use the results obtained by the students as reported in the worksheets. You may also point out that, although it was not brought out in the above discussion (purposefully, though mentioned on page 31 of the module), the face-centered cubic structure and the cubic closest-packed structure are identical. They differ only in the way they are viewed. This interesting point can be brought out by a pyramid of close-packed atoms that has been dubbed the "Pauling Pyramid."

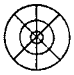
Using 84 polystyrene balls 5 cm to 8 cm in diameter, construct the *Pauling Pyramid* as illustrated in the figures below. For the purpose of clarity, eight of the balls should be a different color, as indicated by the shaded balls.

PAULING PYRAMID

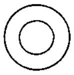
bottom or first layer



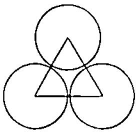
Key:



corner of FCC unit, must be colored



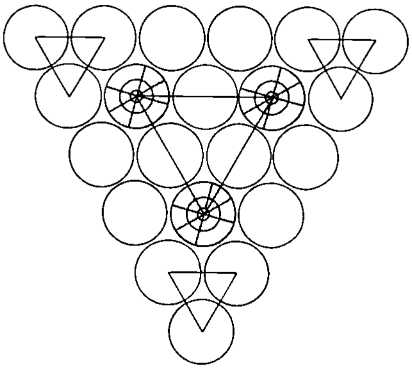
optionally colored



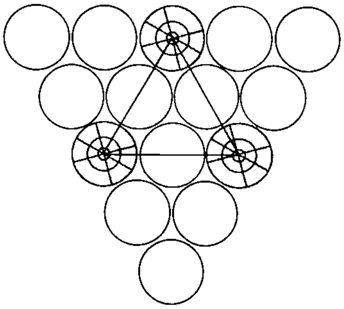
atoms glued together

The base of the pyramid can be held in place by a rack like that used to arrange billiard balls. A more convenient method is to assemble sets of three balls glued together (shown by connecting lines in the diagram). These triangular sets form a solid base and also help hold in the upper layers. Construct the layers as shown in the diagram. The corners of an FCC unit cell are designated by the colored atoms. By carefully removing the other atoms, you can remove this unit cell in one piece. It should look like the figure in the top margin on page 34 of the student module.

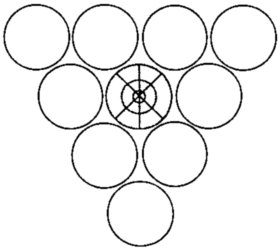
second layer



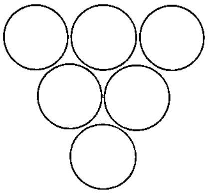
third layer



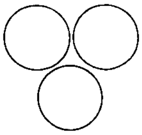
fourth layer



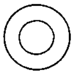
fifth layer



sixth layer



top atom



Ask the class what type of packing is exhibited by the pyramid. Slowly dissect the pyramid, layer by layer, to reveal that the apical atom in the center does not repeat until three layers have been removed, hence an ABC structure, or cubic closest-packing. Continue to dissect until the unit cell is exposed. Remove it and show its resemblance to the figure in the top margin on page 34 of the student module.

The students will more readily see the identity if you have previously shown them a Styrofoam model of an FCC unit cell.

I-15 METALLIC PROPERTIES: LOOK AGAIN

There are only three common lattices for metal atoms in the solid state: body-centered cubic, face-centered cubic (or cubic closest-packed), and hexagonal closest-packed.

FCC (CCP)			*HCP*		*BCC*			
Ca	Fe	Cu	Be	Ti	Li	Ca	Ti	Cr
Sr	Co	Ag	Mg	Hf	Na	Sr	Zr	Mo
Sc	Ni	Au	Ca	Zr	K	Ba	Hf	W
La	Rh		Sr	Mo	Rb		V	Fe
	Ir		Sc	Co	Cs		Nb	U
	Pd		Y	Ni			Ta	
	Pt		La					

Note that HCP structures have an easy glide plane (see page 34 of student module). CCP lattices have even easier glide planes (though we felt they were less easily illustrated). Some of the most malleable metals are Ag and Au (CCP and only one valence electron = "thin molasses"). HCP metals are also usually malleable, except for metals such as Cr, Mo, and W (6 valence electrons = "thick molassess"). BCC lattices have difficult glide planes and tend to be hard. Exceptions are the Group IA and Group IIA metals (1, 2 electrons = "thin molasses"). Note that some metals crystallize in more than one lattice. As expected, for example, CCP-iron is softer than BCC-iron.

All of the above is based on an "ideal" crystal. Some crystal defects that occur are:

1. Vacancies—a lattice point is unoccupied.
2. Interstitials—atoms occupy positions between lattice points.
3. Dislocations—one layer may be incomplete, causing a shift or tilt in the next layer of atoms.
4. Chemical defects—a foreign atom (impurity) is incorporated into the lattice.

The presence of defects can be either good or bad, depending upon the properties that we desire in the metal. Metals often decrease in structural strength as lattice defects increase. On the other hand, the hardness of steel and various alloys can arise from the presence of appropriate amounts of "impurity" elements (e.g., carbon in steel). These impurities alter the properties of the metal. One way that this may be accomplished (and one that is fairly easy to demonstrate visually with the models previously used) is the presence of an atom of a different size at a point in the lattice. Imagine this in terms of the figures on page 34 and the resulting difficulty of gliding in the top figure if a large atom "blocked" the glide. The rate of a chemical reaction may depend upon crystal defects (they provide a "point of attack"). The etching of metal with acids may occur at the defect site.

The presence of a few atoms of arsenic in the borderline metal germanium increases its conductivity (by providing more current-carrying electrons); crystals such as this are called semiconductors and are important in solid-state electronics equipment (e.g., transistors).

Some students might enjoy trying to show crystal defects with the crystal models used in section *I-14*.

Note: The remarkable photographs taken by Professor E. W. Mueller of Pennsylvania State University (student module, pages 25 and 33) show the actual location of atoms in the crystals. Your students may ask why some of the atoms seem to be arranged in ways familiar to them but then seem to change to other ways. The answer is that the photograph represents the point of a pin, and we are seeing sections of several *different* layers, since the tip was shaped to a point.

MINIEXPERIMENT
I-16 GROWTH IS THE NAME OF THE GAME

The purpose of this miniexperiment is to give the student firsthand experience in "growing" ionic crystals.

Concepts

- Ionic crystals exhibit regular lattices.
- Ionic crystals can be very soluble in water.
- Formation of a crystal from a saturated solution can be aided by the presence of solid particles (impurities, or "seeds").

Objective

- "Grow" ionic crystals.

Estimated Time One period to prepare. Observation of crystals may take place in successive days.

Student Grouping Pairs or groups of four

Materials

200–300 cm^3 saturated copper sulfate
200–300 cm^3 saturated potassium aluminum sulfate
200–300 cm^3 saturated ammonium aluminum sulfate
200–300 cm^3 saturated potassium permanganate
60 petri dishes or petri dish covers
15 Bunsen burners
15 wire gauze, asbestos centers

Advance Preparation To save time and materials you may prepare the saturated solutions ahead of time and provide them as stock solutions. Make a saturated solution by dissolving as much salt (see below) as possible in a clean beaker (500-cm^3) about two-thirds full of distilled water. The solution process will be more rapid if the solution is warmed gently. When no more salt appears to dissolve, add a small excess, stir, and allow to stand overnight. The next day, if the excess salt has dissolved, saturation was not reached; add more salt and repeat.

The solubilities of these salts (in grams per 100 cm^3 cold water) are as follows:

copper sulfate ($CuSO_4 \cdot 5H_2O$)	31.6
potassium aluminum sulfate [$KAl(SO_4)_2 \cdot 12H_2O$]	11.4
ammonium aluminum sulfate [$NH_4Al(SO_4)_2 \cdot 12H_2O$]	15.0
potassium permanganate ($KMnO_4$)	2.8

You can also make seed crystals by following the above steps. To get a larger selection of crystals, allow the solution to evaporate until nearly dry (several days). Pick out a nicely formed seed crystal.

Laboratory Tips The purpose of bringing the solution to a boil is to evaporate a portion of the water and thus make the solution supersaturated when it cools. If the student heats the solution too briskly, the burner may have to be cut off before all of the solution has been thoroughly heated, or, the solution may boil over with a gush, making a mess. If you are careful, you can have students return their excess solution decanted from crystals and thus save on materials.

Miniexperiment If you would like to use another lab activity earlier in this module to break up some of the stretches of discussion and nonlab experiments, you can try to grow large crystals, either individually or as a class project.

In order to grow large, regular crystals, three things are important: (1) The growing *solution must be saturated;* (2) a good *seed crystal must be used;* (3) *dust and dirt must be kept out,* since they provide nucleating sites for little crystals to start. In addition, if the temperature in the laboratory fluctuates too greatly, uniform growth will not result. You can stabilize the effect of fluctuating temperatures by using a large water bath (a bucket of water will do) as a temperature buffer.

Potassium aluminum sulfate [$KAl(SO_4)_2 \cdot 12H_2O$], or "alum," and potassium chromium sulfate [$KCr(SO_4)_2 \cdot 12\ H_2O$], or "chrome alum," are very easy to grow. Alum forms clear octahedra, and chrome alum yields purple octahedra. Mixed crystals of the above two salts can be grown in any proportion, since the two compounds are isomorphous. This results in pink, lavender, or light purple crystals. In addition, a perfectly clear and colorless layer of alum can be grown over a purple crystal of chrome alum by changing the solution surrounding the crystal.

There are several other materials that form interesting crystals, such as "Rochelle Salt," or potassium sodium tartrate; nickel sulfate ($NiSO_4 \cdot 6H_2O$ and $NiSO_4 \cdot 7H_2O$); copper sulfate ($CuSO_4 \cdot 5H_2O$). All of these require somewhat more skill, patience, and careful procedure. We recommend *Crystals and Crystal Growing,* by Alan Holden and Phylis Singer (New York: Doubleday, 1960) for excellent directions on growing crystals. If you really want to get into growing crystals in a big way, you may be interested in "The Poor Chemist's Rotary Crystallizer," by L. Kershnar and M. P. Goodstein, which appeared in *Chemistry* (March 1977), p. 25.

Students enjoy observing microscopic crystal growth. Ask the biology teacher to lend you a few microscopes (you might wipe each objective lens with a damp tissue before returning them), spread a small fraction of a drop of NaCl solution on a microscope slide. Using 100X magnification, focus in on the edge of the solution and watch for crystal growth. Do not concentrate too long on crystals already formed, but watch the region where crystals are growing. Growth rate may be accelerated by warming the slide. Adjust the mirror or light source to get either a "dark field" or a "light field" for some interesting observations. NaCl cubes are unmistakable; however, remind the students that the crystals are three-dimensional and they must use the microscope to focus through them. Other solutions that may be tried are: lead acetate, copper sulfate, potassium sulfate, potassium dichromate, potassium chloride, ammonium chloride, sodium bromide, and sodium iodide.

I-17 GETTING A CHARGE

Even though the study of ionic crystal lattices makes sections *I-17* through *I-20* similar to the previous ones, it must be pointed out to the student that the situation here is quite different because we are now dealing with charged particles. Not only does the size of the particles affect the coordination numbers (as in the metals), but so do the electrostatic attractions and repulsions. This is brought out in the following miniexperiment.

MINIEXPERIMENT
I-18 ATTRACTING OPPOSITES

The purpose of this miniexperiment is to show the students an adequate model for the electrostatic attractions and repulsions involved in ionic packing.

Concepts

- Particles of like charge repel each other, and particles of unlike charge attract each other.
- Any physical structure involving charged particles will be influenced by the interaction of attractive and repulsive forces.

Objective

- Build models of ionic packing systems, using disc magnets to represent the ionic species.

Estimated Time 10 to 15 minutes

Student Grouping Pairs or groups of four (depending on the number of disc magnets available)

Materials

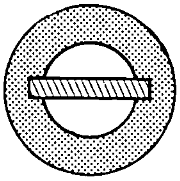
100 disc magnets
30 plates of glass or clear plastic, 8 × 8 cm (optional)
1 overhead projector and screen (optional)
10 ring magnets (optional)

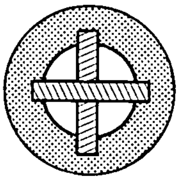
Advance Preparations Paint the magnets red and blue, or black and white, or "+" and "−" to label their polarity. To insure correct coloring of the same polarity, you will find it helpful to stack up the disc magnets. They then will be lined up according to polarity, and all you have to do is successively paint the one on top.

Laboratory Tips Although it looks easy to arrange six "negative" magnets around a "positive" one as shown in the illustration in the student module (Figure A, page 36), it is impossible. The repulsion from the two magnets of similar "charge" on each side of the sixth magnet repels it more strongly than the central magnet attracts it. This is not immediately apparent to the student, however, since the magnets tend to "hop up" on one another, creating the impression that, if you could just keep them all lying flat, the arrangement would work. You can show dramatically that it will not work by fixing one magnet between two parallel plates of glass or clear plastic with shims to allow just a little space for movement. (To force the magnets to remain planar—i.e., not to "hop up"—keep the shims as thin as possible, consistent with elimination of binding by the plates.)

This arrangement holds the magnets coplanar, and, upon pushing the sixth magnet in, one of the other five is forcibly ejected. If you use an overhead projector, students will get a kick out of watching attempts to squeeze in the sixth magnet.

After students are convinced that a coordination number of six in a single plane won't work, they can start experimenting, trying to build a stable lattice with their magnets. Many will come up with a simple square lattice, and this leads directly into the next experiment. You can show this on an overhead projector, too, if you get some ring magnets with holes in the middle. These can be marked by pasting strips of paper across the hole so that +'s and −'s are formed.





It is a good idea to paint the magnets for your own reference, so that you don't turn one upside down and change its polarity!

Postlab Discussion It should be clear to your students that the packing of ions (charged particles) is strongly influenced by the forces of attraction and repulsion.

EXPERIMENT
I-19 PACKING IONS

The purpose of this experiment is to build a model for the packing structures of sodium chloride (NaCl).

Concept

- The structures of ionic crystals are determined by the attractive and repulsive forces of ions.

Objectives

- Construct three-dimensional models of ionic lattices, using plastic balls and boxes.
- Determine the coordination number of ionic structures.
- Calculate the efficiency of packing of ionic structures.

Estimated Time One-half period

Student Grouping Pairs

Materials

15 plastic boxes (see *I-14*)
15 sets of 150 to 200 polystyrene balls, 1″ (25 mm) diameter (Paint half the balls a different color.)

Advance Preparation The format of this experiment is very similar to that of experiment *I-14*, and all the general instructions found there are applicable here. In other words, for the experiment to be successful, it will help a great deal if the teacher has all the transparencies prepared ahead of time and conducts a running discussion as the exercise progresses.

Prelab Discussion The postlab discussion of *I-18* can serve as prelab for *I-19*.

Postlab Discussion The simple cubic lattice (a good model for the NaCl structure) has a coordination number of 6 (six anions about each cation, and six cations about each anion) and a 52.4 percent packing efficiency. The coordination number of the BCC lattice (the model for CsCl) is 8 (eight anions about each cation, and eight cations about each anion), and its efficiency is 67.8 percent.

Note that although the CsCl lattice (Figure 10, student module page 38) is geometrically related to the BCC lattice (Figure 8, page 32), CsCl does *not* have a body-centered lattice. This is a common error.

Differences in ionic radii and their effects on the lattice adopted by an ionic compound have not been discussed in the student module because the topic is probably somewhat too involved for the general high-school course. Further, it is difficult or impossible to conduct packing experiments of the sort done here by using different-sized ions.

For more advanced classes, this topic can be explored somewhat at this time. Models of CsCl and NaCl, made from Styrofoam balls of two sizes (like Figures 10 and 12, student module page 38) and permanently glued together, are useful and easily made (check experiment *I-14*). The relation between size and coordination number is illustrated on page 27 of the module; you can illustrate it further with a modification of the overhead projection of ring magnets (though now of different sizes) described under *Laboratory Tips* for section *I-18*.

For reference, the ionic radii of some common ions (in angstroms) are:

Li^+ = 0.60	Be^{2+} = 0.31	O^{2-} = 1.40	F^- = 1.36
Na^+ = 0.95	Mg^{2+} = 0.65	S^{2-} = 1.84	Cl^- = 1.81
K^+ = 1.33	Ca^{2+} = 0.99	Se^{2-} = 1.98	Br^- = 1.95
Rb^+ = 1.48	Sr^{2+} = 1.13	Te^{2-} = 2.21	I^- = 2.16
Cs^+ = 1.69			
Cr^{3+} = 0.62	Mn^{2+} = 0.82	Fe^{2+} = 0.78	Fe^{3+} = 0.65
Co^{2+} = 0.74	Co^{3+} = 0.53	Ni^{2+} = 0.70	Cu^{2+} = 0.73

The chemist uses a quantity called the "radius ratio" to explain and predict the effect of ion size on the stability or crystal lattices. See *The Nature of the Chemical Bond,* by Linus Pauling (Ithaca, New York: Cornell University Press, 1960) or *Inorganic Chemistry: Principles* of *Structure and Reactivity,* 2nd ed., by James E. Huheey (New York: Harper & Row, 1978).

MINIEXPERIMENT
I-20 THE EFFECTS OF HANGING TOGETHER

The purpose of this experiment is to show to the student some of the properties of ionic crystals that are commonly offered as evidence for the existence of ions.

Concept

- Ionic compounds, metals, and nonmetals have different and distinct properties.

Objectives

- Determine the common physical properties of ionic compounds.
- Compare physical properties of metals and nonmetals with those of ionic compounds.

Estimated Time 15 to 20 minutes

Student Grouping Pairs

Materials The materials needed for this miniexperiment will depend on which properties are chosen for observation.

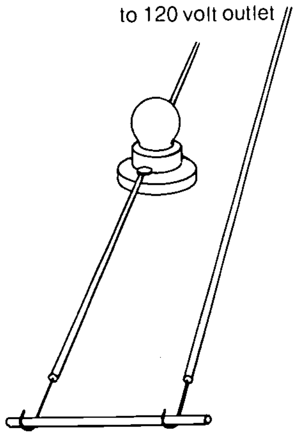
Laboratory Tips To allow the teacher some choice in determining which properties are to be investigated, the following possibilities are suggested:

1. *Appearance.* Examine the specimen for color, appearance, etc. Examples are NaCl, colorless; $CuSO_4 \cdot 5H_2O$, blue; $Co(NO_3)_2 \cdot 6H_2O$, purple; $K_2Cr_2O_7$, orange, etc.
2. *Brittleness.* Ionic crystals many be shattered readily by pounding or, more quietly, by crushing in a mortar. Large crystals of alum $[KAl(SO_4)_2 \cdot 12H_2O]$ or photo "hypo" $(Na_2S_2O_3 \cdot 5H_2O)$ are good.
3. *Solubility.* Ionic crystals often (but not always) dissolve in water but do not dissolve in organic solvents such as toluene. You may simply wish to discuss this without the experiment.
4. *High melting point.* The high melting point of most ionic compounds can be demonstrated relatively easily by heating a little of the compound in a test tube with a Bunsen burner. Some compounds will melt reluctantly and others will not melt at all. Suggested: NaCl, 800°C; KI, 723°C; or other alkali halides. *Do not use hydrates, since the salt may dissolve in its own water of crystallization, giving a false impression of its "melting point."*
5. *Conductivity.* Use the conductivity apparatus from miniexperiment *I-12.* The student can show that
   - **a.** solid ionic compounds are nonconductors,
   - **b.** solutions of ionic compounds in water *do* conduct electricity.

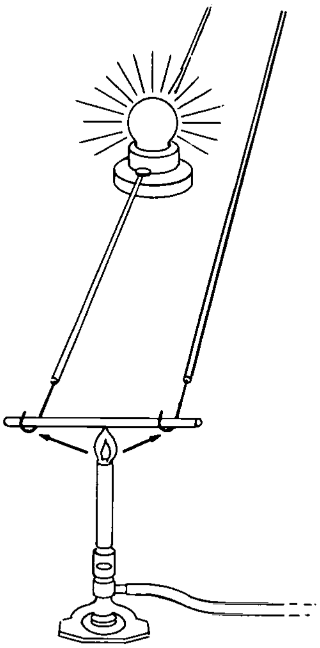
   Further, the teacher can demonstrate the conductivity of molten salts by melting an ionic compound with a relatively low melting point. Suggested: $NaNO_3$, m.p. = 307°C; KNO_3, m.p. = 334°C; LiI, m.p. = 446°C; the last has the advantage of being a simple two-element compound, though it requires higher temperatures. **Caution:** *Molten nitrates are very powerful oxidizing agents. Do not allow them to come in contact with any organic matter!*

   - **c. Caution:** *Conduct this experiment as a demonstration.* Soft glass contains enough ions to conduct current when molten. You can show how heated glass conducts electricity by using soft (lime, flint) glass tubing as a section in a current containing a light bulb (see *illustration*). The unheated glass tube breaks the electric circuit, and the light bulb remains unlit. Heat the glass tube until the glass begins to flow. The light bulb then will glow. As temperature increases, conductivity increases, since the glass becomes mobile, thus allowing ions ease of movement.

GLASS AS A CONDUCTOR



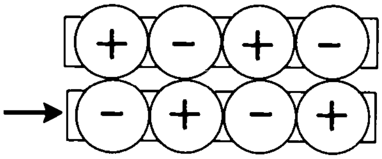
broken circuit
glass rod unheated



complete circuit
glass rod heated

Postlab Discussion In the discussion following the miniexperiment, in which the properties of ionic compounds are interpreted in terms of their structures and the attraction of oppositely charged ions, good use can be made of magnets on an overhead projector (see teacher's guide *I-18*). The mechanism for the splitting of crystals discussed in *I-18* can be dramatically illustrated by means of magnets. Glue the magnets together in two strips (otherwise they will tend to shift as you move them).

Displacing the strip relative to the other by the length of one magnet will cause the two strips to fly apart. The use of disc magnets with center holes and the "charge" marked "+" or "−" with paper strips (see teacher's guide *I-18*) is advantageous if done on an overhead projector.



Inorganic Molecules

The discussion of inorganic molecules shifts the study toward the chemical phenomena emphasized in the remainder of the module. Covalent molecules of the nonmetals make up a large part of this unit through a discussion of their chemical and physical properties. These properties are explained in terms of molecular shape and the effects nonbonding electron pairs may have on molecular geometry. Ideas about partial electronic charge and polar molecules lead to the concept of hydrogen bonding. The use of balloons, magnets, and molecular models makes experiment *I-23* a concrete learning experience.

Transition elements present in catalysts and enzymes are introduced through the role they play in the synthesis of inorganic molecules such as ammonia. Finally, algae growth and pollution are related to phosphates in the environment.

I-21 THE COVALENT BOND: SHARE OR ELSE!

This section starts with a review of the stable octet and covalent bonding studied previously in *Reactions and Reason: An Introductory Chemistry Module*. The octet rule does not explain bonding. It must be looked on as a generalization arrived at by observation of bond formation. Even the hydrogen atom follows the pattern that the octet rule conveys. Two hydrogen atoms are energetically more stable when bonded together than when they exist as separated hydrogen atoms. Since the total capacity of a hydrogen atom is two electrons, it is following the generalization stated in the octet rule when it forms a single covalent bond. This is the typical behavior of hydrogen found in almost all of its compounds.

The rare atypical behavior of hydrogen is its formation of an ionic bond with one of the active metals. In a reaction with lithium, for instance, under the proper conditions hydrogen will accept an electron and become the negatively charged hydride ion.

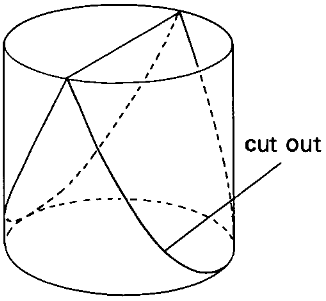
I-22 MOLECULAR SHAPE

The following demonstration can be used to emphasize the need for using three-dimensional models to illustrate molecular geometry. Once the structure is visualized, other properties start becoming evident.

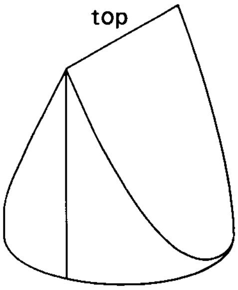
Demonstration *Zeus:* A simple demonstration will help you to reinforce your students' understanding of isomers. The demonstration also emphasizes the importance of observation, problem solving, and model making.

Needed for the demonstration is an object we will call *Zeus* (and henceforth you may wish to refer to this demonstration as "Zeus"). The object may be constructed of wood, Styrofoam, cork, or any material easily carved and painted or colored. Select a convenient dimension (3 to 8 cm) that we will call *d*. The dimension will depend on the materials you have to work with. Make a cylinder whose diameter and height are both *d* units long. Draw a line across the face on one end of the cylinder, dividing the face into halves. Begin cutting at the line on an angle so that the cut terminates at the outer margin of the bottom circle. See the figure here:

1.

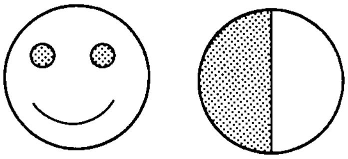


2.



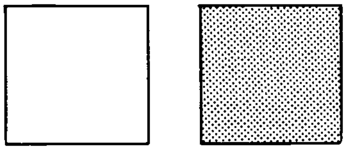
Paint eyes and mouth on the full circle face and paint each of the beveled sides a different color, so the figure will appear like these as shown here when viewed from different angles.

two views of circle



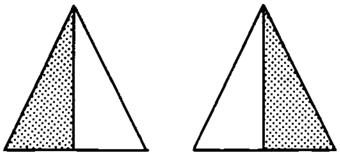
bottom view top view

two views of square



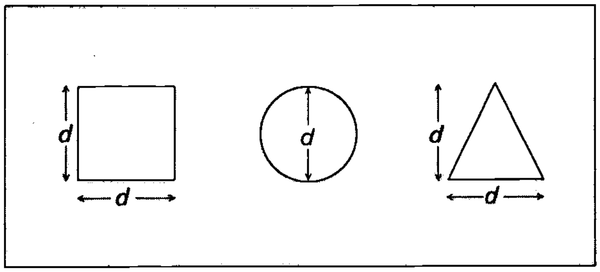
side 1 view side 2 view

two views of triangle



front side view back side view

Now make a template from a piece of cardboard or manila folder by cutting out a circle, a square, and a triangle (not equilateral) of the dimensions shown here.



You can now use Zeus and the template in the creation of a three-dimensional model. Show the students the template and challenge them to construct a solid object that will pass through each of the three holes, completely filling each hole at some point in passing through it. Give your students a chance to design their own Zeus. Then, if necessary, show them how Zeus passes through each hole in the template. In this instance, Zeus serves as a model that meets desired specifications.

The use of models is also sometimes necessary to illustrate in some concrete way something that is either too abstract or is of such a nature that it would not be possible to use the real thing. For example, we cannot bring the solar system into the classroom, so we use a model instead. In chemistry, it is important to have models to help understand the structure of individual molecules. Some of the most basic concepts in chemistry depend on one's understanding of molecular structure. Molecules may be drawn easily on the board, but we can better use models to illustrate the complex structures of many compounds.

If we place our Zeus model on a surface at eye level, and position four students at compass point directions around the model, so that each is facing the object and sees only a single surface, we have established in terms of the object four different points of reference.

One student sees a red square; another sees a white square; a third sees a red-and-white triangle; and a fourth, a white-and-red triangle. If we were to place the object on a clear surface, another student beneath it would view a single-color circle, while a person viewing the object from above would see a half-white and half-red circle. Zeus is also helpful in demonstrating another "difficult to see" concept in chemistry, namely that of isomerism.

During the observations of Zeus by the students, there were two different kinds of triangles observed. If students do not realize that they saw very different triangles, have them reexamine the models by looking at the triangles and the remainder of the model in relation to the triangle. Hint that the two triangles are mirror images of each other. Provide a hand mirror to prove it.

Now, if the concept has been developed to this point, by inductive reasoning the next step should also follow —that the student can now construct molecular models of isomeric compounds. Get out the molecular models and get started. Keep Zeus handy, though. He will enjoy the show and besides, one should not cast early idols so easily aside.

EXPERIMENT
I-23 STRUCTURES OF MOLECULES

The purpose of this experiment is to investigate the geometry of some common molecules.

Concepts

- Molecules have different geometric shapes.
- The shape of a molecule is determined by the bonding and nonbonding pairs of electrons available.

Objectives

- Write dot structures for molecules.
- Write line formulas for molecules.
- Predict the shapes of molecules, including those with nonbonding pairs of electrons.
- Construct ball-and-stick models of molecules.
- Write correct formulas for some compounds.
- Use magnets as models to determine molecular geometry.
- Use balloons as models to determine molecular geometry.

Estimated Time One and one-half to two periods

Student Grouping Pairs or groups of four

Materials

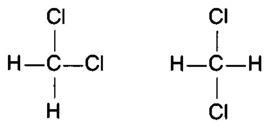
100 disc magnets (See *I-18*)
150 round balloons
15 ball-and-stick model kits (or an adequate substitute: toothpicks and gumdrops, marshmallows, or balls of clay)

Advance Preparation If substitutes are used for the model kits, the teacher must supply additional instructions regarding the bond angles and directions.

You may wish to use specifically prepared inexpensive kits to accommodate *Part 2* of this experiment. Lab-Aids Inc., 130 Wilbur Place, Bohemia, New York 11716 produces several student kits with enough materials for up to 50 students. The atoms in each kit consist of plastic nuclei with pegs set at correct bonding angles to produce accurate three-dimensional models. The kits contain separate procedures, worksheets, and a teacher's guide prepared by the company. Write the company for price list and description.

Prelab Discussion The preliminary discussion should call attention to the number and arrangement of electrons in the atoms as they are found in the periodic table. In this experiment only those elements called "representative" are investigated. No element beyond the second period is considered. Later sections of this module will speak of properties and behavior of transition elements. Note that the first element in a period has one electron. Each succeeding element has one more electron than its preceding neighbor. All of the molecules considered in the experiment are characterized by covalent bonds. Emphasize that covalent bonds arise from electrons being shared. Stress that pairs of electrons will be found as far from each other as possible, and stable molecules have all electrons paired.

Laboratory Tips *Part 1: Graphic Formulas.* Start this discussion by showing the class an example of how dot structures and line structures are drawn. As students work their dot structures and line structures, they should see the difficulty in trying to describe three dimensions. As an exercise, ask them to consider any differences between the two line structures:



$H{-}\overset{Cl}{\underset{H}{C}}{-}Cl \qquad H{-}\overset{Cl}{\underset{Cl}{C}}{-}H$

The fact that the above two structures represent the same molecule may be used to introduce *Part 2* or may be left for discussion at the conclusion of the exercise.

Part 2: Molecular Models. If model kits are limited, the class may be broken into three groups: "carbon group," "nitrogen group," and "oxygen group."

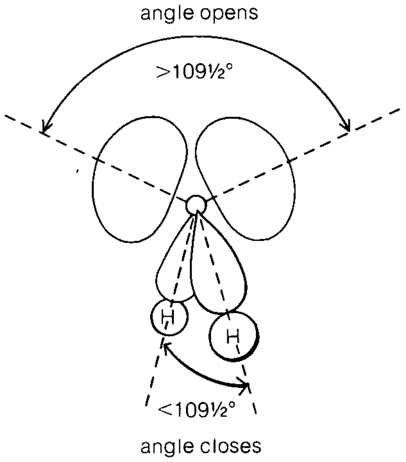
However, pedagogically it may prove more advantageous for students to start by constructing H_2, CH_4, NH_3, and H_2O, in that order. They will thus see in a more dramatic way the need to take into consideration the nonbonding pairs of ammonia and water.

Part 3: Using Magnets to Determine Molecular Geometry. This part shows graphically the repulsion between like charges. Although magnets must necessarily be used on a two-dimensional plane, students should be able to visualize how the like charges could be even more widely separated if allowed to spread over three dimensions. This idea is reinforced by *Part 4.*

Part 4: Using Balloons to Determine Molecular Geometry. Notice that the balloons fill all space around the interlocking necks. The inflated balloons (be sure balloons are rigidly inflated and equal in size) are to be visualized as pairs of electrons. The nucleus that they surround is invisible but is located at the center, where the necks of the balloons are joined.

With this aid, students should be able to visualize pairs of electrons in space around an atom; consequently they should "see" how the number of electron pairs around the atom and the number of bonds that it forms give it a characteristic geometry. However, the experiment does not account for differences in bond angles for molecules that have the same number of electron pairs around the central atom (e.g., $CH_4 = 109.5°$, $NH_3 = 107°$, $H_2O = 104.5°$). Challenge inquiring students to try to find balloons with different shapes corresponding to bonding and nonbonding electron pairs. The student who is successful in this will be able to reproduce qualitatively the geometry of H_2O, for example.

WATER MOLECULE GEOMETRY



The following discussion may help illustrate the differences arising from bonding vs. nonbonding electron pairs:*

In a molecule containing four bonding pairs (e.g., CH_4) or four nonbonding pairs (the Ne atom), all pairs of electrons will be equivalent and occupy the same amount of space. The methane molecule is thus perfectly tetrahedral and the four electron pairs in neon are presumably also arranged in a perfect tetrahedral arrangement (this is not experimentally verifiable). In a molecule such as water, however, the four pairs are no longer equivalent. Two pairs have protons (hydrogen nuclei) situated in such a way that they attract electrons towards them.

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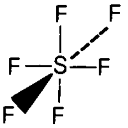
*For further discussion, see B. E. Douglas and D. H. McDaniel, *Concepts and Models of Inorganic Chemistry* (New York: John Wiley & Sons, 1965), pp. 61–62, and J. E. Huheey, *Inorganic Chemistry: Principles of Structure and Reactivity* 2nd ed., (New York: Harper & Row, Publishers, 1978), pp. 197–210.

The electron clouds of these pairs become somewhat elongated and "thinner" than the clouds of the nonbonded pairs that are attracted *only* by the central (oxygen) nucleus. The latter orbitals will tend to "spread out" and occupy more space about the oxygen than the "thinner" bonding pairs.

It should be pointed out that although the balloon experiment may seem to be a little elementary, it truthfully reflects the behavior of electron pairs as discussed above. The valence-shell electron-pair repulsion (VSEPR) theory is currently much used in inorganic chemistry. Some students might like to read Peter Dahl's article, "Valence-Shell Electron-Pair Repulsion Theory," in *Chemistry* (March 1973), pages 17–19. You might read R. J. Gillespie's articles on the VSEPR theory in *Journal of Chemical Education* (January 1963), page 295; (January 1970), pages 18–23; and (June 1974), pages 367–370.

Use the questions asked in *Part 4* and the results obtained by your class. Ask the students to derive rules for predicting structures of molecules. Reference can be made to drawings of various molecules in the module. Please note that the students may not be familiar with the convention

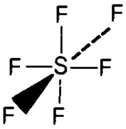


indicating that the lower left bond is extending *toward* the reader and the dashed bond in the upper right is extending *away* from the reader. Refer the student to the drawings of space-filling models. It is good at this point to reinforce the concept introduced in *Reactions and Reason: An Introductory Chemistry Module* that models are only convenient *representatives* of reality. Note, for example, that although atoms are spherical, the models representing fluorine atoms (student module page 42) are not spherical, because they have to fit together to form the fluorine molecule, the model of which, again, is only approximately representative of the real molecule. This distinction is an important one and easily forgotten.

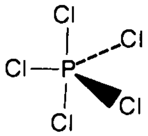
ANSWERS TO PROBLEMS

(Student Module page 49)

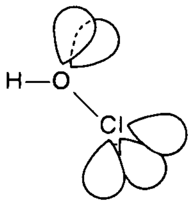
1. Sulfur hexafluoride (SF_6)



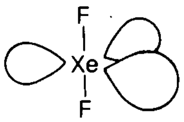
2. Phosphorus pentachloride (PCl_5)



3. Hypochlorous acid (HOCl)



4. Xenon difluoride (XeF_2)

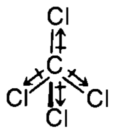


I-24 A MATTER OF LIFE AND BREATH

The geometry of the carbon dioxide molecule is linear, whereas water is bent. This would make little difference if electron pairs were shared equally between bonded atoms. In these two molecules—carbon dioxide and water—oxygen pulls the shared electron pairs closer to itself and consequently farther away from the atom to which it is bonded. This gives oxygen a small negative charge and leaves the opposite atom with a small positive charge. This is called bond polarity. The polarity of certain molecules is an observable fact. A dipole moment may result if the molecule is bent or has a nonsymmetrical shape. Then the small negative charge on one side of the molecule and the small positive charge on the other side will create a polar molecule.

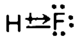
The polarity can be looked at as the vector sum of the polarities of the bonds in the molecule. The dipole moment, μ, is used as a measure of the polarity of the molecule and is measured in Debye units (D). Following are some examples of molecules with polar bonds.

Since the chlorine atoms are symmetrically distributed around the carbon, the vector sum of bond polarities is zero. Carbon tetrachloride is a nonpolar molecule.



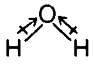
$\mu = O$

The bond polarity between the two atoms that make up hydrogen fluoride serve here as our example of the simplest of polar molecules.



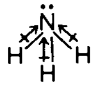
$\mu = 1.75\ D$

The water molecule exists in a linear or bent form with a bond angle close to 105°. The vectors that result from the two polar bonds gives an appreciable dipole moment to this molecule.



$\mu = 1.84.D$

The bond angles in an ammonia molecule are reduced to 107° by the nonbonding pair of electrons. The result of the three polar bonds gives the ammonia molecule an appreciable dipole moment.



$\mu = 1.46\ D$

Further development along these lines may be found in many textbooks.*

ANSWER TO PROBLEM

(Student module page 53)

Ice floats because it has a lower density than liquid water. This may be stated a different way: a given quantity (mass) of water has a greater volume in the form of ice than in the liquid state. Thus, water is one of the few liquids that expands as it freezes, with possibly disastrous effects on automobile radiators.

I-25 SOME IMPORTANT MOLECULES

Demonstration A simple demonstration of the action of a catalyst is as follows. Place a few cm^3 of methanol in an Erlenmeyer flask and swirl the methanol around enough to fill the flask with vapor. Heat a platinum wire briefly in a burner flame and insert it in the flask (but not into the liquid). The wire will continue to glow from the heat given off by the reaction

$$2CH_3OH + O_2 \rightarrow 2CH_2O + 2H_2O + \text{heat}$$

(This is the same type of reaction that takes place in the small flameless pocket heaters.) The wire will continue to glow for a long time if it is allowed to hang in the flask. If the reaction is allowed to proceed for any length of time, put the flask in a hood, since the formaldehyde formed is irritating. Copper wire can also be used as a catalyst, but it is less effective than platinum.

The subject of nitrogen fixation is one of much current interest among chemists working in the area of bioinorganic chemistry. There is the pragmatic aspect: in a world desperately in need of increasing food production to match increasing population, nitrogenous fertilizers are greatly in demand. At the more basic scientific level, scientists are intrigued by the question of how the enzyme nitrogenase works in living systems. The illustrations on page 55 of the student module portray the nitrogen-fixating structures in legumes at several levels of organization: (1) the entire legume, red clover (upper right), and Austrian peas (upper left); (2) the bacteria-containing root-nodules on clover and soybeans (left); and (3) an electron micrograph of the actual bacteria in the nodule that are responsible for the nitrogen fixation (lower right).

For a discussion of the current state of research on nitrogen fixation, see "Biological Nitrogen Fixation," by W. J. Brill, *Scientific American* (March 1977), pages 68–81, and J. E. Huheey, *Inorganic Chemistry: Principles of Structure and Reactivity,* 2nd ed., (1978), pages 758–762.

ANSWERS TO PROBLEMS

(Student module page 57)

1. Na and Fe are metals; H, S, O, N, and Cl are nonmetals.
2. (a) H_2O; (b) NH_3; (c) NaCl; (d) H_2SO_4; (e) HCl
3. Na^+Cl^- is ionic; lattice is shown in Figure 12, page 38. The H_2O and NH_3 molecules are discussed in section *I-23 Laboratory Tips;* H_2SO_4 consists of tetrahedral sulfur with two =O and two —OH groups on it.

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*See R. T. Morrison and R. H. Boyd, *Organic Chemistry*, 3rd ed. (Boston: Allyn & Bacon, 1973), p. 22 ff.

4. Water, ammonia, and sulfuric acid are discussed in this section. (These are the compounds anticipated in this question; HCl will be encountered in the next section as an acid cleaner; NaCl is present in blood, perspiration, and sea water—it is also important as a melting agent for ice and as a source of Na and Cl.)

EVALUATION ITEMS

These are evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. As the sizes of the atoms in a metallic lattice increase,

   A. the coordination number increases.
   B. the efficiency of packing decreases.
   C. the density increases.
   [✓ D.] none of the above necessarily happens.

2. The coordination number of the hexagonal closest-packed (HCP) structure is:

   A. 6 B. 8 C. 10 [✓ D.] 12

3. The coordination number of the simple cubic structure is:

   A. 4 [✓ B.] 6 C. 8 D. 12

4. A box 12 × 16 × 6 cm is filled with marbles 2 cm in diameter. Using a certain arrangement, 155 marbles are packed. The space occupied by the marbles in the box is 12 × 16 × 5 cm. The volume of the sphere is given by $V = 1.33 \times 3.14 \times r^3$. The efficiency of this packing is:

   A. 52% B. 82% [✓ C.] 68% D. 74%

5. Which of the following are the correct coordination number and efficiency of the cubic closest-packed (CCP) structure?

   A. 6; 52% B. 8; 68% C. 12; 68% [✓ D.] 12; 74%

6. Which of the following statements is false?

   A. Conductivity is one of the distinctive characteristics of metals.
   [✓ B.] The highly ordered structure of a simple cubic lattice is favored by many metals.
   C. The malleability of metals is related to the presence of easy gliding planes in the lattice.
   D. The cubic closest-packed structure and the face-centered cubic structure are identical.

7. The coordination number of the NaCl ionic structure is:

   A. 1 B. 4 [✓ C.] 6 D. 8

8. Which of the statements below is true?

   A. With a few exceptions, most ionic solids are good electrical conductors.
   B. The highly ordered structure of a simple cubic lattice is favored by many metals.
   [✓ C.] Most ionic solids have high melting points resulting from the electrical forces holding them together.
   D. The hexagonal closest-packed (HCP) structure and the face-centered cubic (FCC) structure are identical.

9. Consider the following compounds:

HCl	H_2O	$CaCl_2$	CH_4	H_2	NiF_2
I	II	III	IV	V	VI

   a. Pure covalent bonding is found in:

   A. I B. II C. III [✓ D.] V

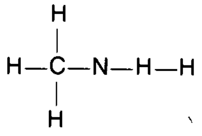
   b. Polar covalent bonding is found in:

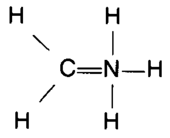
   [✓ A.] II B. III C. V D. VI

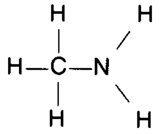
10. Which of the following molecules has a nonbonding pair of electrons?

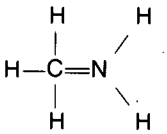
   A. PCl_5 B. CH_4 C. SF_6 [✓ D.] none of the above

11. The correct line structure for methylamine, CH_3NH_2 is:

A.


C.


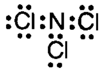
[✓ B.]


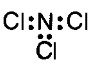
D.


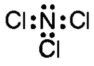
12. Which of the following molecules has a linear structure?

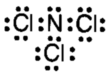
   [✓ A.] CO_2 B. H_2O C. NH_3 D. CH_4

13. The correct electron-dot structure for NCl_3 is:

A.


C.


B.


[✓ D.]


Acids and Bases

Since many of the inorganic molecules discussed previously react with water to produce acids or bases, it is convenient to consider these substances at this time. The initial definitions of acids and bases presented are those used to explain reactions occurring in aqueous solution. Hence an acid is an H^+ donor, a base is an H^+ acceptor, and water is the neutralization product. Relevancy is introduced through an experiment in which acid or basic substances common in the home are used in titrations. The concluding section introduces the Lewis acid-base concept. A base is an electron-pair donor and an acid is an electron-pair acceptor. Thus, positive metal ions, Lewis acids, offer a convenient way to introduce the next section, *Chemistry of the Transition Elements*.

Demonstrations may be used to illustrate the acid and base colors of the indicators noted in the module. Natural dyes may be included. Try aqueous solutions or the ethanol extracts of red cabbage, the red skins of radishes, or dark berries such as blueberries, fresh, frozen, or even in jellies and jams.

I-26 GETTING DOWN TO BASICS (AND ACIDICS)

Emphasis is placed on the idea that acids and bases are "opposites" and thus neutralize each other, regardless of the definition used. Indicators used to determine endpoints in neutralization titrations are organic compounds that are weak acids or bases. The indicators change color because they represent an equilibrium situation that is affected by H_3O^+ or OH^- concentrations. A discussion of pH has been omitted but may be included if desired.

The pH scale is a scale that indicates the relative acidity or basicity of a solution compared with water. Values usually range from 0–14, with any value less than 7 indicating an acid and any value above 7 indicating a base.

EXPERIMENT
I-27 HOUSEHOLD ACIDS AND BASES

Concept

- Acids and bases neutralize each other quantitatively.

Objectives

- Perform an acid-base titration to the correct indicator endpoint.
- Calculate the percentage of acid or base in household cleaners.

Estimated Time Two and one-half periods

Student Grouping Pairs

Materials

30 50-cm^3 burets
45 250-cm^3 Erlenmeyer flasks
phenolphthalein indicator, 1 percent
3 liters 1 M NaOH
3 liters 1 M HCl
vinegar
household cleaners (liquid Vanish, Lysol, liquid Saniflush, solid Saniflush, Drano)
15 double buret clamps
15 75-mm diameter funnels
1 tube stopcock grease (depending on burets used)

Advance Preparation Three liters 1 M NaOH can be prepared by dissolving 120 g NaOH pellets to make 3 liters of solution. Three liters 1 M HCl can be prepared by diluting 250 cm^3 concentrated HCl to 3 liters of solution.

If a more accurate standard is desired, the NaOH can be titrated against a primary standard such as KHPh (potassium acid phthalate) and then the HCl solution is titrated versus the newly standardized NaOH. More conveniently, 0.1N HCl and 0.1N NaOH can be purchased and used as standards.

You may wish to work up sample problem sheets to illustrate the calculations involved in this experiment.

Prelab Discussion Demonstrate the use of a buret. Emphasize the need for accuracy in recording the exact endpoint and avoiding parallax in making buret readings. Remind students to swirl the titration vessel constantly. If this is the students' first titration, some time and effort should be spent on practicing techniques.

Discuss the role and use of indicators. Review the concept of molarity and relate it to solution concentrations. It is suggested that a sample calculation be performed before the experiment.

Caution: *Point out to the students that the cleaners involved in this experiment are strong acids and bases and should be handled accordingly.*

Laboratory Tips For the first part of the experiment, everyone titrates vinegar with NaOH. For the second part, try to get an even distribution of acidic and basic household cleaners.

Titration of Vanish will be difficult, since the color of the indicator tends to fade. The manufacturers of Vanish are currently adding 0.15% potassium monopersulfate, and this oxidizing agent destroys the indicator. It may be restored by adding 1 drop of 0.1 M $KMnO_4$ solution.

Acidic cleaners that fizz do so because manufacturers sometimes add sodium bicarbonate to make the user think that something "is happening" when, of course, all that is happening is that some of the acid present is wasted! These cleaners will give low results. This could be used as part of a postlab discussion. For more on this, see R. Breedlove, "The Chemistry of Toilet Bowl Cleaners," *Chemistry* (November 1973), p. 31.

Postlab Discussion Compare results with manufacturers' claims. Here are some reasonable results.

Vinegar	5% CH_3COOH
Lysol	8.5% HCl
Liquid Vanish	9.25% HCl
Liquid Saniflush	7% HCl and 2% $H_2C_2O_4$
Solid Saniflush	81% $NaHSO_4$
Solid Vanish	75% $NaHSO_4$
Bowlene	83% $NaHSO_4$
Commercial ammonia	28% NH_3
Drāno	54.2% NaOH

Note that the manufacturers may change their formulations, so check the label on the container. If you wish to run this experiment as an "unknown," you can place opaque tape over the label or transfer the contents to a flask before presenting the cleaner to the students.

ANSWERS TO PROBLEMS

(Student module page 64)

1. 11.7% HCl
2. $Fe(OH)_3 + 3HCl \longrightarrow Fe^{3+} + 3H_2O + 3Cl^-$

I-28 ACIDS AND BASES GO TO WORK

You may wish to amplify this discussion by looking up representative industrial processes involving acids and bases. You can then relate them to everyday chemistry. Note that acids resulting from sulfur pollutants in the atmosphere are discussed on page 87 of the student module and in *Form and Function: An Organic Chemistry Module* and in *The Delicate Balance: An Energy and the Environment Chemistry Module*. Sulfuric acid is a good example. Sulfur dioxide, obtained from roasting copper ores, for many years was allowed to escape into the atmosphere where "sulfuric acid production" took place.

$$2Cu_2S + 3O_2 \longrightarrow 2Cu_2O + 2SO_2$$
$$2Cu_2O + Cu_2S \longrightarrow 6Cu + SO_2$$

More recently, use of oxygen-enriched air has resulted in high-grade sulfur dioxide that can be converted into sulfuric acid (see student module page 56):

$$2SO_2 + O_2 \longrightarrow 2SO_3$$

Although sulfur trioxide will combine with water, as shown in the module, the reaction is slow, and normally sulfuric acid is used to dissolve both reactants:

$SO_3 + H_2SO_4 \longrightarrow H_2S_2O_7$ (disulfuric acid, pyrosulfuric acid, or "oleum") $H_2O + H_2S_2O_7 \longrightarrow 2H_2SO_4$

The reaction of sulfur dioxide with oxygen and water in the atmosphere (together with a similar reaction for nitrogen oxides) seems to be the main culprit in the phenomenon known as "acid rain" or "acid precipitation." The pH of heavily industrialized areas in North America and Europe has been measured as lower than 5.6 (the expected pH value for pure water in equilibrium with ambient concentrations of atmospheric carbon dioxide). For more on this, see G. E. Likens, *Chemical and Engineering News* (November 22, 1976).

About half of the sulfuric acid produced industrially is used to make fertilizers; ammonium sulfate can be made directly from sulfuric acid and ammonia (see pages 53–57 for the latter):

$$H_2SO_4 + 2NH_3 \longrightarrow (NH_4)_2SO_4$$

Insoluble phosphate rock can be solubilized by sulfuric acid:

$$Ca_3(PO_4)_2 + 2H_2SO_4 \longrightarrow \underbrace{Ca(H_2PO_4)_2 + 2\,CaSO_4}_{\text{"superphosphate"}}$$

These two simple processes provide fertilizers containing the important nutrients: nitrogen, potassium, and sulfur.

The high boiling point of sulfuric acid makes it a convenient reagent for the preparation of volatile acids:

$$NaCl + H_2SO_4 \longrightarrow HCl(g) + NaHSO_4$$

Because it is cheap, sulfuric acid is often the acid of choice if one simply needs a source of hydrogen ions in an industrial process. The iron and steel industries use sulfuric acid solutions ("pickling baths") to remove oxide scale:

$$Fe_2O_3 + 3H_2SO_4 \longrightarrow Fe_2(SO_4)_3 + 3H_2O$$

Hot concentrated sulfuric acid is a good oxidizing agent, typified by reactions such as:

$$Cu + 2H_2SO_4 \longrightarrow CuSO_4 + 2H_2O + SO_2$$

Large quantities of sulfuric acid are used in the petroleum and related industries to oxidize tars and similar impurities so that they may be removed.

I-29 WHAT'S IN A NAME?

Lewis acid-base definitions are introduced to illustrate additional ways of defining acids and bases. Emphasize that acids and bases are still opposites in this definition. The octet rule is used as the basis for Lewis structures.

There is a difference between accepting a *pair* of electrons from a Lewis base to form a *covalent bond* and the complete transfer of an electron to form an *ionic* bond. It is useful to distinguish between them. Only the first type of reaction is included in the Lewis concept. If desired, students may practice writing reactions of this type, using Lewis dot structures:

$$A \left\{ \begin{array}{l} NH_3 + BCl_3 \longrightarrow H_3NBCl_3 \\ PH_3 + BF_3 \longrightarrow H_3PBF_3 \\ NH_3 + H^+ \longrightarrow NH_4^+ \end{array} \right.$$

$$B \left\{ \begin{array}{l} CO + BF_3 \longrightarrow OCBF_3 \\ Ag^+ + 4NH_3 \longrightarrow [Ag(NH_3)_4]^+ \\ Cu^{2+} + 4NH_3 \longrightarrow [Cu(NH_3)_4]^{2+} \end{array} \right.$$

Reactions of Type A are very similar to that shown in the student module and should present few problems. Type B reactions are somewhat more complicated and may require more discussion (CO has two lone pairs; which does it use?, etc.).

$OC{:}BF_3$ $[Ag(NH_3)_4{}^+]$ $[Cu(NH_3)_4{}^{2+}]$

Note: If you are more familiar with $Ag(NH_3)_2{}^+$ you may have wondered about the above; the tetraammine complex forms in excess concentrated ammonia.

ANSWERS TO PROBLEMS

(Student module page 66)

1. HCl, H_2SO_4, and BCl_3 are acids. Both HCl and H_2SO_4 are H^+-donors; BCl_3 is a Lewis acid.
   $NaOH$, NH_3, and KOH are bases.
2. $2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$
   $3HCl + Al(OH)_3 \longrightarrow AlCl_3 + 3H_2O$
   $H_3PO_4 + 3NH_3 \longrightarrow (NH_4)_3PO_4$
   or
   $3NH_3 + 3H_2O \longrightarrow 3NH_4{}^+ + 3OH^-$
   and
   $3NH_4{}^+ + 3OH^- + H_3PO_4 \longrightarrow$
   $3H_2O + 3NH_4{}^+ + PO_4{}^{3-} \longrightarrow (NH_4)_3PO_4$

EVALUATION ITEMS

These are evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. H_2O can act as either an acid or a base. The reaction that best illustrates the behavior of water as a base is:

   [✓ A.] $HCl + H_2O \longrightarrow H_3O^+ + Cl^-$
   B. $H_2O + NH_3 \longrightarrow NH_4{}^+ + OH^-$
   C. $HCl + NaOH \longrightarrow NaCl + H_2O$
   D. $H_2O + NH_2{}^- \longrightarrow NH_3 + OH^-$

2. In a titration of vinegar the following data were gathered:

Mass of empty flask:	32.5 g
Mass of flask and vinegar:	73.5 g
Volume of 1.0 M NaOH used:	33.5 cm^3

   The percentage of acetic acid (molar mass 60) in this vinegar sample is:

   A. 4.5% [✓ B.] 4.9% C. 5.3% D. 5.7%

3. A commercial household cleaner contains KOH (molar mass 56) as its active ingredient. To determine its strength, a titration was run and the following results were obtained:

Mass of cleaner:	6.3 g
Volume of 1.0 M HCl used:	24.8 cm^3

The percentage of active KOH in the cleaner is:

A. 18% B. 20% C. 22% D. 24%

4. Two samples of household cleaners were titrated for their ammonia content. The results are as follows:

100 cm^3 of cleaner A required 75 cm^3 of 0.15 M HCl to neutralize.

100 cm^3 of cleaner B required 50 cm^3 of 0.20 M HCl to neutralize.

If both cleaners cost the same amount per pint, which is the best buy?

A. cleaner B C. both are equal
B. cleaner A D. cannot tell from data provided

5. Which of the following is a Lewis acid?

A. BF_3 B. OH^- C. $CH_3CO_2^-$ D. NH_3

Chemistry of the Transition Elements

Let us consider certain characteristics of the transition elements—their beauty and color, for instance, and oxidation and reduction concepts. Ligands acting as Lewis bases are discussed according to their formation of coordination compounds. A series of laboratory syntheses of complex compounds permits the student to enjoy using some new skills. Complex compounds, their colors, and their structures are considered, along with applications. Coordination compounds are used in experiments in water softening and solvent extraction. We close with an experiment in which a complexing agent gives a specific test for the presence of lead, which may or may not be present in samples brought in by the students.

I-30 YOU CAN'T LOSE FOR WINNING

Redox reactions are introduced, using traditional definitions of oxidation and reduction. The number of electrons lost by transition metals cannot be predicted accurately. The idea of multiple oxidation numbers stresses the wide range of reactions possible for transition metals.

Plan to show the CHEM Study film, *Vanadium—A Transition Element,* if it is available.

Demonstration *An Inorganic Amoeba:* On a watch glass, place a small pool of mercury (about the diameter of a dime). Add 6 M H_2SO_4 to cover the mercury pool. Add about 1 cm^3 of 0.1 M $K_2Cr_2O_7$. ($KMnO_4$ may also be used—just enough to make the solution pink.)

In the presence of electrolytes the surface tension of mercury decreases and the pool of metal flattens out. Touch it on the outer edge with a piece of iron wire (nail or needle) and it suddenly contracts. Fix the wire in such a way that the contraction breaks the contact. The mercury will flatten again; touch the wire, it will contract again; and so on. This rhythmic pulsing of the mercury can be shown to the class by using an overhead projector. Place a few bits of organic shrubbery about the watch glass and chide the students about the life of this "inorganic amoeba." Finally, emphasize that this is a redox reaction by discussing the equations for the reaction, which are approximately as follows.

First, $K_2Cr_2O_7$ and $KMnO_4$ are strong *oxidizing agents*—that is, they contain metals in high oxidation states that are seeking electrons. The mercury is able to furnish electrons

$$Hg \longrightarrow Hg^+ + e^-$$

meanwhile being oxidized itself. You may notice a grayish film, probably mercury(I) sulfate, forming on the surface of the mercury. The pool of mercury becomes relatively electron-poor. Touching the mercury pool with the iron wire (an even better *reducing agent* or source of electrons than mercury) reduces the mercury compound (the film disappears) and the iron transfers excess electrons to the mercury pool. The iron is, of course, oxidized in the process:

$$Fe \longrightarrow Fe^{2+} + 2e^-$$

When the iron loses contact with the mercury, the process repeats itself. As to the relation between electric charge and surface tension, we suggest you consult your physics teacher.

The overall reactions are, therefore, reduction of dichromate or permanganate:

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O$$
$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$$

and oxidation of the iron wire:

$$Fe \longrightarrow Fe^{2+} + 2e^-$$

Fe^{2+} can, of course, be oxidized to Fe^{3+} by either of these oxidizing agents, but as long as there is metallic iron present it will tend to produce only +2 ions because of the reaction:

$$2Fe^{3+} + Fe \longrightarrow 3Fe^{2+}$$

The mercury cycles back and forth; in a sense it acts as a catalyst promoting the redox reaction. If students ask why the oxidizing agent doesn't react directly with the iron wire, you may answer that it probably does, though slowly, and that there is plenty of iron and oxidizing agent to allow them to react with the mercury as well. If you care to pursue if further, however, you might explain that chromate (or dichromate) tends to form a protective film on iron. If the radiator protector you add after a flush is yellowish, chances are it contains chromate. You may find that this seemingly simple demonstration leads to a very involved discussion. Carry it only as far as the interest and ability of your students warrant.

Although this experiment has been conducted many times, we are not aware that anyone has ever completely characterized the products. See J. A. Campbell, *Journal of Chemical Education* 34 (1957), page 362, or *Tested Demonstrations in Chemistry,* 6th ed. (Easton, Penna.: Chemical Education Publishing Company, 1965), pages 138 and 143.

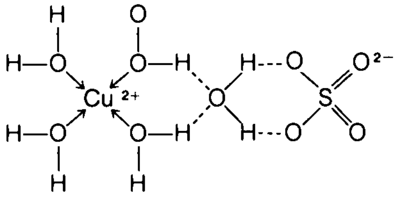
I-31 COORDINATION CHEMISTRY! HOLD ON!

The term *ligand* is new to the student. Point out that all ligands are not identical, but all act identically; that is, all ligands must be able to donate a pair of electrons. They are all Lewis bases.

The removal and addition of ligands are readily demonstrated, using copper(II) sulfate. Place a few crystals of $CuSO_4 \cdot 5H_2O$ into a test tube. Hold horizontally and heat over a Bunsen burner.

$$\underset{\text{hydrated-blue}}{CuSO_4 \cdot 5H_2O} \longrightarrow \underset{\text{anhydrous-white}}{CuSO_4} + 5H_2O$$

Water will condense on the wall of the test tube, leaving behind the white anhydrous $CuSO_4$. After this is cool the addition of water to the anhydrous $CuSO_4$ will cause heat to be evolved, with a return to the blue color that characterizes the hydrated form.



electron pairs are denoted by ⟶

Four water molecules donate electron pairs to the copper ion. The fifth molecule of water is held between the tetraaquacopper(II) ion and the sulfate ion by hydrogen bonds and need not be of much concern.

Note: Avoid excessive heating, indicated by a darkening of the copper sulfate sample and an evolution of irritating SO_3 fumes.

$$CuSO_4 \longrightarrow \underset{\text{black}}{CuO} + \underset{\text{white fumes}}{SO_3}$$

EXPERIMENT
I-32 GET COORDINATED!

The purpose of the experiment is twofold:

1. To acquaint the student with the synthesis of new compounds, an area where much of the work of practicing inorganic chemists is done.
2. To provide the class with a nice variety of coordination compounds from which it may draw some generalizations about this type of compound.

Concepts

- Lewis bases may form coordination compounds with metallic ions.

- The properties of coordination compounds are dependent on the way the metallic ions and the ligands share the electrons.

Objective

- Synthesize, isolate, and purify coordination compounds.

Estimated Time Two to three periods (depending on the number of syntheses done)

Student Grouping Pairs

Materials It is not expected that all students will carry out all of the syntheses. One or two for each pair of students should prove adequate. A judicious assignment on the part of the teacher will insure a good sampling of all the coordination compounds for discussion purposes.

Refer to *Advance Preparation* for specific quantities of chemicals for each compound being synthesized. Refer to the student module for specific apparatus requirements for each synthesis.

50-cm^3 beakers
150-cm^3 beakers
250-cm^3 beakers
600-cm^3 beakers
125-cm^3 Erlenmeyer flasks
250-cm^3 Erlenmeyer flasks
10-cm^3 graduated cylinders
50-cm^3 graduated cylinders
Büchner funnels and adapters
250-cm^3 filter flasks
rubber tubing
filter paper, medium weight
ice
Bunsen burners
hot plate (optional)
aspirators
30-cm^3 vials
glass tubing, 6-mm diameter
drying oven (optional)
watch glasses
ring stands
rings
thermometers −10°C to 110°C
wire gauze, asbestos centers

Advance Preparation You will need the following quantities of chemicals for each pair of students doing a given synthesis. The amount of chemicals and apparatus required depends upon the number of students doing a given synthesis.

Part 1: Some Complexes of Copper(II)

A. Tetraamminecopper(II) sulfate:
12.5 g $CuSO_4 \cdot 5H_2O$
30 cm^3 conc. NH_3
30 cm^3 ethanol
10 cm^3 acetone

B. Bis(ethylenediamine)copper(II) sulfate:
12.5 g $CuSO_4 \cdot 5H_2O$
17 cm^3 ethylenediamine
40 cm^3 ethanol
10 cm^3 acetone

Part 2: Some Complexes of Cobalt(III)

A. Hexaamminecobalt(III) chloride:
1 g powdered charcoal
10 g NH_4Cl
15 g $CoCl_2 \cdot 6H_2O$
30 cm^3 conc. NH_3
30 cm^3 30% H_2O_2
25 cm^3 conc. HCl
25 cm^3 ethanol
15 cm^3 acetone

B. Sodium hexanitrocobaltate(III):
30 g $NaNO_2$
10 g $Co(NO_3)_2 \cdot 6H_2O$
18 cm^3 9 M $HC_2H_3O_2$
70 cm^3 ethanol
10 cm^3 acetone

Part 3: Some Complexes of the Ligand Acetylacetone

A. Cobalt(III) acetylacetonate:
2.5 g $CoCO_3$
20 cm^3 acetylacetone (2,4-pentanedione)
30 cm^3 10% H_2O_2

B. Chromium(III) acetylacetonate:
2.7 g $CrCl_3 \cdot 6H_2O$
20 g urea
6 cm^3 acetylacetone (2,4-pentanedione)

C. Aluminum acetylacetonate:
6 cm^3 acetylacetone (2,4-pentanedione)
6 g $Al_2(SO_4)_3 \cdot 18H_2O$
30 cm^3 6 M NH_3 solution
litmus paper, red and blue

D. Iron(III) acetylacetonate:
$3\ g\ FeCl_3 \cdot 6H_2O$
$6\ cm^3$ acetylacetone (2,4-pentanedione)
$30\ cm^3\ 6\ M\ NH_3$ solution
litmus paper, red and blue

Laboratory Tips The procedures in these syntheses are time-consuming. The usual class period makes it necessary to introduce shortcuts in order for students to make their complex compounds in a reasonable amount of time. Heating, cooling, and allowing time for crystallizing have been reduced. These shortcuts reduce yield and purity of some compounds. The students should be informed that these are shortcuts and not standard procedures.

In spite of shortcut measures, some of these syntheses require more than one lab period. We have tried to make them fit two successive days.

Part 1: Some Complexes of Copper(II)

A. Tetraamminecopper(II) sulfate:

The standard preparation for this compound suggests letting the solution stand for an hour before chilling. This would improve crystal size and total yield, but we feel a 15-minute wait is adequate. Tetraamminecopper(II) sulfate is deep blue. Long, needlelike crystals of this compound can be formed by adding the ethyl alcohol to the blue solution carefully so that it floats on top. Stopper and set the flask aside, without mixing, for about a week. The alcohol slowly diffuses into the solution and reduces the solubility of the solute. The slow growth tends to produce larger crystals.

B. Bis(ethylenediamine)copper(II) sulfate:

Bis(ethylenediamine)copper(II) sulfate is deep blue.

Part 2: Some Complexes of Cobalt(III)

A. Hexaamminecobalt(III) chloride:

The standard preparation calls for two hours in the water bath. We have obtained good results with the present procedure. By the time the experiment is resumed the next day, it will be at room temperature, allowing the student to filter immediately. Hexaamminecobalt(III) chloride is orange-yellow.

B. Sodium hexanitrocobaltate(III):

This experiment has been shortened to save time. The standard preparation involves a 30-minute oxidation and an hour for crystallization. Sodium hexanitrocobaltate(III) is brown-orange.

The directions for the synthesis of sodium hexanitrocobaltate(III) call for an air current to be bubbled through the reaction flask. Since most laboratories do not have an air source available, a suitable one can be rigged up using a two-hole stopper and bent glass tubing connected to an aspirator. The suction created will draw air bubbles through the solution, thus enabling the NO produced to escape.

Part 3: Some Complexes of the ligand acetylacetone

A. Cobalt(III) acetylacetonate:

If no oven is available, the crystals may be air-dried overnight. The crystals may be recrystallized, if desired, from $50\ cm^3$ of hot toluene to which is added $300\ cm^3$ of heptane or petroleum ether. The mixture is then cooled and filtered. **Caution:** *These solvents are flammable!* Cobalt(III) acetylacetonate is dark green.

B. Chromium(III) acetylacetonate:

If a steam bath is not available, the beaker may be set in a hot water bath on a hot plate overnight, but be sure the temperature control is set low enough so that the solution does not evaporate. Chromium(III) acetylacetonate is purple.

C. Aluminum acetylacetonate:

If desired, the product may be recrystallized by dissolving in the least possible quantity of toluene and reprecipitating it by the addition of petroleum ether. **Caution:** *These solvents are flammable!* Separate on a Büchner funnel and wash once with petroleum ether. Dry in air. Aluminum acetylacetonate is white.

D. Iron(III) acetylacetonate:

Iron(III) acetylacetonate is red-brown. The last three syntheses are very simple—merely mix each one and filter off—although the chromium compound requires an overnight reaction. All three can be done by one person without difficulty in two periods.

Postlab Discussion You may wish to collect the samples in labeled vials prepared by the students and organize these for use with section *I-33*, which serves as the ongoing postlab for this experiment.

I-33 LOOK TO THE RAINBOW

Common properties of coordination compounds are discussed with little recourse to experiment because no experimental proof is possible at this level. One property that can be gone into in some detail is the color of the complexes, one of their most striking features. In order to let the students come to their own conclusions, we have kept the conclusions in the module to a minimum. This means that in order for this to be effective, you must lead a vigorous discussion, elucidating and challenging suggestions and ideas from students. Discussing *Answers to Problems* is an excellent place to start.

ANSWERS TO PROBLEMS

(Student module pages 77–78)

1. The ligand alone does not determine the color:

   a. $Cu(NH_3)_4SO_4$ = blue
   $Co(NH_3)_6Cl_3$ = yellow
   b. $Cu(en)_2SO_4$ = blue
   $Co(en)_3Cl_3$ = yellow
   c. $Co(acac)_3$ = green
   $Cr(acac)_3$ = purple
   $Al(acac)_3$ = white
   $Fe(acac)_3$ = red-brown
   No.

2. Although there is some relation between the metal ion and the color of a complex, the colors can vary:

   a. $Cu(NH_3)_4SO_4$ = blue
   $Cu(en)_2SO_4$ = blue
   b. $Co(NH_3)_6Cl_3$ = orange-yellow
   $Na_3Co(NO_2)_6$ = brown-orange
   $Co(en)_3Cl_3$ = yellow
   $Co(acac)_3$ = green
   No.

The following is a short outline of suggestions for a lively review and discussion. Note: The first two questions involve further discussion of two of the problems on pages 77–78 of the student module.

1. *Does the ligand determine the color?* The two ammonia complexes are different colors: blue and yellow. There are two* ethylenediamine complexes, which are also different: blue and yellow. All four acetylacetonate complexes are different colors: green, white, purple, red-brown. The ligand *alone* does *not* determine the color.

2. *Does the metal ion determine the color?* The two Cu^{2+} complexes are deep blue; three of the four Co^{3+} complexes are some shade of orange. Clearly, there does seem to be some relation between the color of a complex and the metal ion present but, as shown by $Co(acac)_3$, the colors can vary.

3. *What are some further factors?* From here on, the discussion can turn any one of several ways and so no attempt has been made to outline it in the module. In the discussion, students may come up with some of the following observations:

- The original copper sulfate solution was blue, though not so intense as the two complexes formed. Point out that this is a result of a complex formed with water molecules: $Cu(H_2O)_4^{2+}$.
- Ask about the colors formed when Cl^- (from NH_4Cl or HCl) was added in *Part 2A*. The deep blue color is from the $CoCl_4^{2-}$ complex. Ask why that complex is blue but three other cobalt complexes were orange. Try to get someone to suggest (a) the oxidation state is different, Co^{2+} *vs.* Co^{3+}; (b) there are four ligands instead of six. (Both answers are correct; both factors play an important role in color.) You can conclude that the metal ion must be in a particular oxidation state [you may wish to provide further evidence of this in the form of a sample of copper(I) halide for comparison] and have the same number of ligands.
- Pursue the different color of the $Co(acac)_3$ complex as compared with the other three. All four contain Co^{3+} and all have a coordination number of 6. The difference is that the orange complexes have N-Co bonds. Evidently, the nature of the ligand has some effect.

Pursue the above discussion as long as class interest and time warrant. You may have students who are not willing to drop it at this point and who will ask why one is yellow and not blue and why not purple for $Al(acac)_3$, etc. These questions

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*In the first edition, we included the preparation of $[Co(en)_3]$, but the simplified procedure that was given for student use was not reliable—results were erratic. (The correct preparation is not difficult, but it is inappropriate to a short working period. You might consider having one student prepare a sample for class discussion. See *Inorganic Synthesis* 2:221; 9:162.)

are difficult to answer with satisfaction without more theory. If *s*, *p*, *d*, and *f* electrons have been discussed, you *can* point out that color is associated with metal ions having one or more *d* electrons—that is, most of the transition metal ions. In contrast, Al^{3+} forms colorless compounds.

Students are often interested in color, and you can further this interest by discussing the relation between the color of a compound and its spectrum. A compound may be blue because the complementary color, red, has been absorbed. Yellow compounds absorb blue and purple light.

Ask your physics teacher to provide you with a suitable prism and a very bright source of light (such as an arc light). With this equipment you can show quite nicely how transition metal compounds absorb colors. Adjust the lamp and prism until a complete spectrum (red, orange, yellow, green, blue, violet) is projected onto a screen or wall. Then place *flatsided cells** (beakers will *not* work—they bend the light rays) containing solutions of $CuSO_4$, $Co(NO_3)_2$, etc., in the path of the beam. In the case of the blue $Cu(H_2O)_4^{2+}$ complex, or the even more intensely colored $Cu(NH_3)_4^{2+}$ complex,** it will be noted that the red end of the spectrum disappears. The solution is blue because it absorbs the reds and oranges, etc.; white light minus red light equals blue light. This is not so much chemistry as merely the idea of complementary colors, but the students may be interested in actually seeing the spectrum and how the red light is absorbed by the solution. It is also possible to buy a kit that allows the same demonstration to be performed with a 35-mm projector. Although we have used it with some success, we find the arc light–prism arrangement more convenient. (For example, how do you explain to your supervisor why there is copper sulfate dripping out the bottom of the Carousel projector?!)

I-34 USING COORDINATION COMPOUNDS

Coordination chemistry is so integral a part of inorganic chemistry that it is almost impossible to overestimate its value. In the sections that follow, the student is introduced to some aspects of the uses of coordination chemistry. We here present you with a few ideas on coordination compounds that you may find interesting and useful when discussing them.

First, it should be remembered that almost all aqueous chemistry of metal ions involves coordination compounds. Chemists Fred Basolo and Ralph Pearson pointed out several years ago: "It is generally agreed that an aqueous ion such as iron(II) is in reality $Fe(H_2O)_6^{2+}$, or something similar. . . . Thus, it follows that coordination chemistry includes the greater part of all inorganic chemistry. . . . It can also be said that knowledge of complex ions is important in many other areas, for example, biochemistry, analytical chemistry, catalysis, electrochemistry, mineralogy, and radiochemistry . . ." (see Basolo and Pearson, *Mechanisms of Inorganic Reactions*, 2nd ed. [New York: John Wiley & Sons, 1967]).

One of the most noticeable properties of complexes of transition metals is their color. From ancient times, we have found uses for intensely colored coordination compounds as pigments. Minerals such as hematite (Fe_2O_3, red), orpiment (As_2S_3, yellow), malachite ($CuCO_3 \cdot Cu(OH)_2$, green), and galena (PbS, black) were crushed for use in paints, among other things. Since they had no really satisfactory blue pigment, the Egyptians invented a process for the manufacture of blue glass from silica, malachite, calcium carbonate, and sodium carbonate. They also manufactured red lead (Pb_3O_4) quite early. Later, pigments containing cyano complexes of Fe(II) and Fe(III) (Prussian blue and Turnbull's blue) were employed. More recently, organic dyes have replaced many inorganic pigments, especially in dyeing cloth, but even here inorganic complexes (mordants) may still be important.

Coordination compounds of various kinds are employed as catalysts. Their ability to accept or donate electrons or to accept or donate ligands allows them to catalyze reactions involving electron transfer or chemical changes of ligands. Biological examples are called enzymes and are discussed later in this module. It may be noted in passing that the poisoning of enzymes discussed on pages 93–95 of the student module involves special affinities of some ligands for certain metals; for example, mercury is strongly bound by

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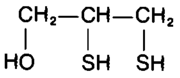
*These are available in some overhead projection kits or you can make some from Plexiglas in the same way as the boxes in *I-14*.

**In aqueous solution both of these complexes probably have two more molecules of water attached, but there is little point in adding unnecessary diversionary information for the student.

sulfur-containing ligands, hence when mercury gets into the body it is strongly bound to (and poisons) sulfur-containing enzymes. The chemist fights back by using complexing compounds containing sulfur, such as *British Anti-Lewisite* (BAL)



$$\underset{HO}{CH_2}-\underset{SH}{CH}-\underset{SH}{CH_2}$$

which tend to tie up the mercury and reduce its poisonous capabilities. As a matter of fact, BAL was invented for a similar purpose, to counteract the poisonous effects of "Lewisite", a gas used in World War I. In this compound the poisonous element is arsenic, which also has an affinity for sulfur. For more on medical uses of chelating agents, see A. Magliulio, *Chemistry* (January 1974), p. 25; and *Chemical and Engineering News* (May 2, 1977).

MINIEXPERIMENT
I-35 BATHTUB RINGS AND THINGS

This miniexperiment investigates the effect of hard water on soap. The role that metal ions play in water hardness is also considered.

Concept

- The presence of certain metal ions leads to "low-sudsing" hard water.

Objective

- Determine the effect of different metal ions on the sudsing ability of soap solutions.

Estimated Time One-half period

Student Grouping Pairs

Materials

100 cm^3 0.1 M aluminum nitrate
100 cm^3 0.1 M calcium nitrate
100 cm^3 0.1 M cobalt nitrate
100 cm^3 0.1 M iron(III) nitrate
100 cm^3 0.1 M magnesium nitrate
1000 cm^3 soap solution
90 18 × 150-mm test tubes

Advance Preparation To prepare the above solutions, dissolve:

3.0 g soap in 1000 cm^3 distilled water
3.8 g $Al(NO_3)_3 \cdot 9H_2O$ in 100 cm^3 distilled water
2.4 g $Ca(NO_3)_2 \cdot 4H_2O$ in 100 cm^3 distilled water
2.9 g $Co(NO_3)_2 \cdot 6\ H_2O$ in 100 cm^3 distilled water
4.0 g $Fe(NO_3)_3 \cdot 9H_2O$ in 100 cm^3 distilled water
2.6 g $Mg(NO_3)_2 \cdot 6H_2O$ in 100 cm^3 distilled water

Laboratory Tips It is wise to check the formulas of salts used to prepare solutions; the solutions should be 0.1 M. If you plan to do miniexperiment *I-36*, some of these solutions are required. Look ahead and prepare a sufficient amount.

Be sure that your soap solution is truly a soap solution (e.g., Ivory Flakes) and not a synthetic detergent. The latter is affected by hard water but does not form soap scum. Also, no water softener can be present in the product if the experiment is to work properly. The concentration of the soap solution should be about 3 g per 1000 cm^3, although it is not critical. This produces a solution that is roughly 0.01 M.

In order to shorten the length of time consumed by this experiment, we have employed only those metals that form precipitates with soap and have used distilled water as the only control. If you care to expand the miniexperiment and show that some metal ions are harmless with respect to water hardness, you can provide the class with 0.1 M solutions of ammonium nitrate, sodium nitrate, and potassium nitrate. These will not interfere with the sudsing action of the soap solution.

If you live in a part of the country with very hard water, you can use some tap water as one of the samples. Remember that the concentrations used here are more concentrated than in natural waters, so less soap will be tied up as "scum" with the natural waters.

Be sure your students realize that the six numbered tubes correspond to the six numbered solutions, a different solution in each tube.

Prelab Discussion Review the mechanism of the action of soap (*Form and Function: An Organic Chemistry Module, O-41* to *O-43*, and *Communities of Molecules: A Physical Chemistry Module*, p. 53). Some metal ions render the soap ineffective by tying up the polar end of the soap molecule and causing the soap to precipitate.

$$\underset{\text{sodium stearate}}{2C_{17}H_{35}CO_2^-Na^+} + Ca^{2+} \longrightarrow \underset{\text{calcium stearate}}{(C_{17}H_{35}CO_2)_2Ca} + 2Na^+$$

Using the above equation, discuss the insolubility of the calcium stearate as forming the soap scum.

Postlab Discussion In discussing the result of this experiment, point out that the hard water used here is quite a bit more concentrated than you are apt to find in your water system. This makes the precipitates more visible. If you live in a soft-water area, do not let your students be misled into thinking that natural hard water gives such startling results as their experiment.

MINIEXPERIMENT
I-36 WATER SOFTENING

The purpose of this miniexperiment is to observe the action of water-softening agents upon the ions that cause water hardness.

Concept

- Hard water may be made soft by the removal of metal ions by precipitation or through the formation of coordination complexes.

Objective

- Determine the effectiveness of different chemicals as water softeners.

Estimated Time One-half period

Student Grouping Pairs

Materials

250 cm³ 0.1 M calcium nitrate
250 cm³ 0.1 M magnesium nitrate
100 cm³ 0.1 M iron(III) nitrate
250 cm³ 0.1 M sodium triphosphate
250 cm³ 0.1 M sodium carbonate
250 cm³ 0.1 M sodium tetraborate
250 cm³ 0.1 M NTA (nitrilotriacetic acid, sodium salt)
1000 cm³ soap solution
150 18 × 150-mm test tubes

Advance Preparation To prepare the above solutions, dissolve:

6.0 g $Ca(NO_3)_2 \cdot 4H_2O$ in 250 cm³ distilled water
6.5 g $Mg(NO_3)_2 \cdot 6H_2O$ in 250 cm³ distilled water
4.0 g $Fe(NO_3)_3 \cdot 9H_2O$ in 100 cm³ distilled water
3.0 g soap in 1000 cm³ distilled water
9.2 g $Na_5P_3O_{10}$ in 250 cm³ of water*
2.7 g Na_2CO_3 in 250 cm³ of water
5.1 g $Na_2B_4O_7$ in 250 cm³ of water
6.4 g $N(CH_2COO)_3HNa_2 \cdot H_2O$ in 250 cm³ water
or
6.6 g $N(CH_2COO)_3Na_3 \cdot H_2O$ in 250 cm³ water

Either of the last two may be used as the "NTA" solution. The sodium salts of nitrilotriacetic acid are used because they are more soluble.

Laboratory Tips Refer to the tips in the previous experiment, *I-35*.

Prelab Discussion Introduce this experiment as one application of coordination complexes, sections *I-33* and *I-34*. The metal ion is from the hard water, while the water softeners furnish the ligands. Also, some softeners act by simple, direct precipitation:

$$Na_2CO_3 + Ca^{2+} \longrightarrow CaCO_3(s) + 2Na^+$$

Postlab Discussion In the postlab discussion of water softeners, you might wish to discuss ion exchange resins. These are now used to a great extent in laboratories (the "distilled water" in your lab may be from an ion exchanger, for example). The students may be familiar with them from the small versions available for preparing water (often erroneously called "water *filters*") for steam irons. In a simplified view, ion exchange resins are organic complexers that can tie up both cations and anions and "in exchange" (hence the name) provide H^+ (for the cation absorbed) and OH^- (for the anion absorbed), which combine to form water.

I-37 SOLVENT EXTRACTION: CHEMICAL TWEEZERS

This section discusses the importance of being able to separate one metal ion from another. The procedure discussed and demonstrated in miniexperiment *I-38* is that of solvent extraction. The procedure capitalizes on the preferential solubility of a coordination compound into one of two immiscible layers.

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*Formerly Calgon was an appropriate mixture of phosphate water softeners and could be used here. After the phosphate scare, Calgon became "low phosphate," with 34 percent $Na_5P_3O_{10}$ and 66 percent $Na_2CO_3 \cdot NaHCO_3$. See J. R. Van Wazer, *Phosphorus and Its Compounds: Inorganic Chemistry* (New York: Interscience, 1958), p. 776.

MINIEXPERIMENT
I-38 CATION LIB

The purpose of this miniexperiment is to provide an example of the solubilities coordination complexes have in two immiscible solvents and to illustrate extraction phenomena.

Concept

- Metallic ions can be separated by formation of coordination compounds that have different solubilities in polar and nonpolar solvents.

Objective

- Suggest a method of separating a mixture of coordination compounds by using the technique of solvent extraction.

Estimated Time One-half period

Student Grouping Pairs

Materials

50 cm^3 0.1 M iron(III) nitrate
50 cm^3 0.1 M cobalt(II) nitrate
100 cm^3 hydrochloric acid, conc.
100 cm^3 hydrobromic acid, conc. (see *Advance Preparation*)
100 cm^3 sodium thiocyanate, saturated
250 cm^3 methyl isobutyl ketone (4-methyl-2-pentanone)

Advance Preparation To prepare the above solutions, dissolve:

1.2 g $Fe(NO_3)_3$ in 50 cm^3 of water
1.5 g $Co(NO_3)_2 \cdot 6H_2O$ in 50 cm^3 of water
100 g NaSCN in 100 cm^3 of water

A substitute for hydrobromic acid may be used in this experiment. Mix 50 cm^3 of concentrated HCl with 50 cm^3 of a saturated solution of KBr or NaBr. Upon mixing, some of the salt will precipitate; decant and use the clear solution labeled as HBr.

Prelab Discussion Section *I-37* provides the introduction for this miniexperiment.

Laboratory Tips The saturated NaBr or KBr solution alone will complex Fe^{3+}, but the solution must be strongly acid if extraction with MIBK is expected to give good results. Many extractions with MIBK may be necessary for weak acid solutions—this fact offers an opportunity to extend this study to include pH as a significant factor in extraction phenomena.

If you choose to use concentrated HBr as such, be sure to check the precautions for handling on the container label, and other reference sources.

Range of Results

		Fe^{3+}			Co^{2+}	
Before MIBK	*Cl$^-$*	*Br$^-$*	*SCN$^-$*	*Cl$^-$*	*Br$^-$*	*SCN$^-$*
(Color after adding complexing agent)	faint yellow	yellow-orange	intense red	deep blue	—	blue
After MIBK						
Color of MIBK layer	yellow	yellow-orange	intense red	—	—	blue
Color of water layer	—	—	—	deep blue	—	—

Postlab Discussion The complex ion $(FeBr_4)^-$ extracts readily into MIBK while $(CoBr_4)^{2-}$ remains in the aqueous phase. This difference in solubilities offers a means of separating these ions. You might like to demonstrate the separation of Fe^{3+} and CO^{2+}. The extraction is most effective if several separatory funnels are used.

Prepare a solution containing:

40 cm^3 1.0 M $Co(NO_3)_2$
20 cm^3 0.1 M $FeCl_3$
60 cm^3 HBr, conc. (see *Advance Preparation*)

The solution contains a 20:1 ratio of Co^{2+} to Fe^{3+} because of the intensity of color of the iron complex. Point out that it is possible to extract a small amount of one species in the presence of a very large amount of another.

Add the above solution to a 250-cm^3 separatory funnel together with 80 cm^3 MIBK. Stopper the funnel and shake briskly. Allow the solvents to separate and then carefully drain off the bottom, aqueous layer into a second separatory funnel. Be careful not to let any of the MIBK layer slip through. Place the first funnel in a rack where the class can see it. Now add 80 cm^3 MIBK to the solution in the second funnel and repeat the extraction.

The extractions may be continued several more times. The initial color is very dark from the $(FeBr_4)^-$ ion. After two or three extractions, the pink color of the $[Co(H_2O)_6]^{2+}$ ion shows clearly in the aqueous layer, but it is still possible to extract more iron from it. Even though there is much more CO^{2+} than Fe^{3+} in aqueous solution; the iron is extracted preferentially.

Extraction procedures such as this are often used in analytical chemistry to separate metal ions before their individual analysis. They are also important industrially in extracting metals from mixtures.

I-39 KNOCK, KNOCK! WHAT'S THERE?

This short section is a "prelab" discussion for miniexperiment *I-40*. Use it to get the students interested in the miniexperiment and discuss with them methods for testing for lead. See the *Prelab Discussion* of *I-40* for further ideas.

MINIEXPERIMENT
I-40 GET THE LEAD OUT!

The purpose of this miniexperiment is to illustrate the use of coordination compounds to detect the presence of specific metallic ions.

Concept

- The distinctive colors of metallic complexes may be used to identify qualitatively the presence of a particular metal.

Objective

- Test qualitatively for the presence or absence of lead in a sample.

Estimated Time One-half period

Student Grouping Pairs

Materials

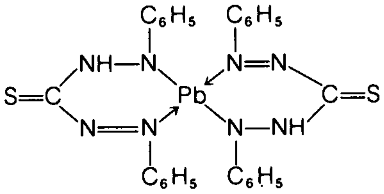
100 cm^3 0.1 M acetic acid
50 cm^3 dithizone solution in TTE (Freon 113 or Du Pont TF Solvent)

Advance Preparation To prepare the 0.1 M acetic acid, dilute 1.14 cm^3 concentrated acetic acid to 200 cm^3.

Dissolve 2 mg of dithizone in 100 cm^3 of TTE (Freon 113). Do not prepare more dithizone than you expect to use in a short time; left standing in a glass bottle, it will extract enough lead from the glass to turn red!

Collecting samples: Samples suspected of containing lead (paint chips, metal seals on wine bottles, metal from toothpaste tubes, or pieces of solder) may be used in the procedure described in the student module.

Prelab Discussion If *The Delicate Balance: An Energy and the Environment Chemistry Module* has not been or will not be taught, concentrate on the usefulness of the test in detecting lead and its importance in the environment. If the module has already been taught, you may wish to concentrate the prelab discussion on the coordination chemistry and solvent extraction involved here. Dithizone forms the complex shown below, which is more soluble in TTE than in water.



Laboratory Tips The procedure given in the student module is modified from *Spot Tests,* by Fritz Feigl (New York: Elsevier, 1954), the classic reference work of its type. We have used the procedure to detect lead in stock solutions, paint chips, solder, and auto exhaust. Acetic acid has been used to solubilize insoluble compounds, but the test is more sensitive when performed in neutral solution. If more sensitive tests are desired, try the following spot test.

A drop of the test solution is spotted on to a small piece of filter paper. A drop of dithizone solution (0.05%) is placed on the filter paper next to the previous drop, so that the two will meet. A reddish color indicates the presence of lead. Placing the sample in a large beaker saturated with NH_3 fumes (from a small amount of concentrated ammonium hydroxide in the bottom) followed by drying in an oven or otherwise makes the test more sensitive. Be sure to run a blank for comparison.

Auto exhaust deposits collected on filter paper from a cold tailpipe can be readily spot tested. Add a couple of drops of glacial acetic acid to the dirtiest smudge of carbon and proceed as above. By turning the filter paper over, you can see the colors unclouded by the spot.

This is an extremely sensitive test for lead, and the importance of running blanks, knowns, and minimizing contamination cannot be overemphasized.

References: A. E. Sherword, *Metallurgia* 80 (1969), p. 209; and Kodak Publication No. JJ-5, *TLC Visualization Reagents and Chromatographic Solvents.*

Postlab Discussion The impact of lead in our environment may generate much discussion. Note that lead poisoning can be extremely serious, especially in children. Refer your students to the article by Dale Jenkins listed on page 101 of their module. Do not let emotionally charged discussions divert attention from the usefulness of this miniexperiment. Stress the use of coordination compounds for specific tests and the extreme sensitivity of these tests. (The dithizone solution in your bottle may be pink—see *Advance Preparation.*)

ANSWERS TO PROBLEMS

(Student module page 85)

1. Any of the equations associated with the syntheses in *I-32* or *I-38* is suitable.
2. Examples discussed in this section include water softening, chemical solvent extraction, and qualitative testing.
3. Although ready formation of coordination compounds and extensive redox behavior (several oxidation states) are emphasized in this section, the student may recall the earlier discussion *(I-7)* where it was pointed out that transition metals are usually less reactive and stronger than representative metals.

EVALUATION ITEMS

These are evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. An oxidation reaction is best represented by:

   A. $CH_3CO_2H + NaOH \longrightarrow CH_3Na + H_2O$
   B. $Cu^{2+} + 4NH_3 \longrightarrow Cu(NH_3)_4^{2+}$
   C. $Fe \longrightarrow Fe^{2+} + 2e^-$
   D. $NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$

2. Which of the following is a reduction reaction?

   A. $Fe^{2+} \longrightarrow Fe^{3+} + e^-$ C. $Fe^{3+} - e^- \longrightarrow Fe^{4+}$
   B. $Fe^{2+} - e^- \longrightarrow Fe^{3+}$ D. $Fe^{3+} + e^- \longrightarrow Fe^{2+}$

3. A ligand commonly encountered in coordination compounds is:

   A. Fe B. Cu^{2+} C. NH_3 D. BF_3

4. The coordination complexes of copper(II) compounds are frequently:

   A. orange B. blue C. white D. red

5. While conducting a series of determinations on hard water, a student ran out of standard soap solution. With no soap available to prepare a new solution, the student substituted a 10 percent detergent solution, made the necessary solution, and continued with the experimental work. Discuss this procedure and its results.

6. A method that could be used to confirm the presence of lead in a sample of gasoline is:

   A. Place 50 cm^3 of gasoline in a dish and let it evaporate; then scrape the lead residue together for a display.
   B. Burn 50 cm^3 of gasoline in a dish; since the lead doesn't burn, it will remain in the bottom of the dish for all to see.
   C. Both A and B.
   D. None of the above.

Bioinorganic Chemistry

We now return to the concept of inorganic compounds as pollutants, but, as students will see, with further study we can balance this "Bad News" with "Good News" by considering essential elements and trace elements necessary to living systems. Metalloenzymes and a schematic explanation of their function, as well as the process of oxygen transport in animals are considered. The concluding section discusses a redox reaction that includes a compound necessary in providing energy for living things—a compound that, fittingly, contains iron, a transition element.

I-41 BIOINORGANICS: THE "BAD NEWS"

The discussion of inorganic pollutants is intended to show that the inorganic chemicals themselves are not harmful, but the way in which they are used can be harmful. Stress the effects of concentration and reaction rates in a qualitative manner.

Cursory reading of newspapers and magazines will turn up a vast supply of articles and references to everyday problems stemming from chemicals in our environment. The latest government decisions concerning the flight of SSTs in the United States, aerosol cans, etc., can lead to a very lively discussion of the ozone problem. It must be stressed that although the problem is far from settled, our uncertainty about the dangers involved suggests perhaps a conservative approach.

The pictures on page 89 of the student module show Lehigh Gap, Pennsylvania, before and after the construction of zinc smelters in the area. The top photo was taken in the 1880s; the lower one, in the 1930s. The area around the smelters in Lehigh Gap is sparsely vegetated or completely barren over an area of about 4 850 000 m^2. The surface soil in the area contains up to 8 percent zinc. See M. J. Jordan, *Ecology* 56 (1975), p. 78.

In *The Delicate Balance: An Energy and the Environment Module,* pollution-related issues and the chemistry involved are discussed in some detail. If your class will not be studying the environment module, or if time allows, you may wish to have them pursue these investigations;

1. If there is a large smelting or refining plant in your community, you might ask students to suggest the kinds of pollutants associated with the manufacturing process and disposal of industrial wastes. What are the effects of these chemicals on human beings? On plant and animal life? Have local practices changed in the handling of pollutants in recent years? Because of the social and economic as well as environmental issues involved, students should be encouraged to consider the numbers of people each plant employs, the monetary costs of correcting pollution, and the health hazard for the workers and the community.
2. If yours is a farming community, or if there are farms nearby, ask students to find out what kinds of insecticides local farmers use, their chemical compositions, how long they've been in use, and ways in which they may be more effective and more or less harmful to the environment than other insecticides.
3. If there are no local examples of possible sources of pollution, perhaps students could research the before-and-after aspects of a major city, such as London or Pittsburgh, which was transformed by a drastic change in laws regulating heating fuels, waste disposal, and other household and industrial practices.

Lead the discussion into section *I-42* and the benefical uses of, and reliance upon, trace metals in life processes.

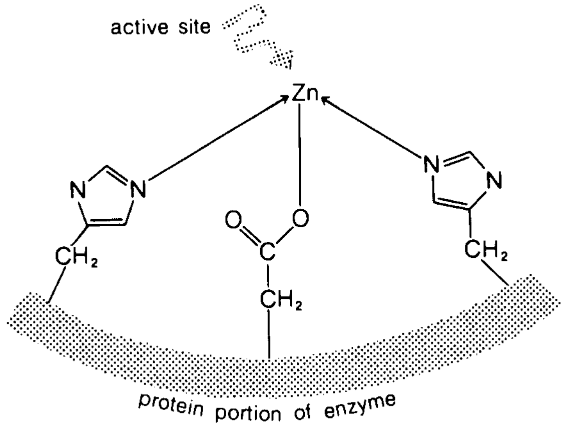
I-42 BIOINORGANICS: THE "GOOD NEWS"

About 20 elements represent the chemicals essential for the life processes. It may be pointed out that two of these elements, hydrogen and oxygen, account for about 89 percent of the atoms in the human body, and carbon and nitrogen for slightly less than 11 percent. This means that the rest, less than 1 percent, must be spread out among the others. They are considered trace elements, since they are needed in very small quantities—some of them so small that their discovery took place only in recent years as more sophisticated analytical techniques were developed. Among the less common trace elements we find: cobalt (an integral component of vitamin B_{12}); zinc (needed for growth and function of blood cells, kidneys, liver, and other organs); chromium (needed for insulin activity); and manganese and molybdenum (required for normal body growth and metabolism).

I-43 ENZYMES: CHEMICAL MEAT CLEAVERS

The remainder of this section can be looked upon as a further example of trace metal activity and a closer look at the mechanism through which the metals operate.

The first example is a metalloenzyme called carboxypeptidase A. The zinc in this enzyme uses its coordination bonding to grip and hold peptide chains in position so that the enzymes can then separate amino acids from the chain.



Enzymes are a very important class of compounds and are discussed not only here, but also on pages 54–56 of this module and at some length in *Molecules in Living Systems: A Biochemistry Module*. They are useful in portraying various features of chemistry: catalysis, reaction rates, "poisoning" of both biological and chemical systems, and the importance of stereochemistry and molecular shape.

I-44 TAKE A DEEP BREATH!

Without the complexing ability of the iron in hemoglobin and myoglobin, the blood would be unable to carry sufficient oxygen to the cells of the body. The mechanism by which this is accomplished appears, at first sight, to be an exceedingly simple one: Hemoglobin picks up oxygen from the lungs and transfers it through the bloodstream to myoglobin in the tissues, which holds it until the cells need it to oxidize food and supply energy. The resultant carbon dioxide dissolves in the blood and is released in the lungs, completing the cycle. However, on closer inspection, a number of questions arise: Why does hemoglobin transfer the O_2 to the myoglobin? This implies that the latter has a stronger affinity, but why? Since tissues have more CO_2 and a lower pH (remember, carbonic acid is a weak acid) than the lungs, does pH have any effect on hemoglobin? (The answer is "yes"; it's called the *Bohr effect.*)

As one might expect from a molecule of molecular mass of about 64 500, the mechanisms involved when oxygen coordinates to hemoglobin are rather complex but finely tuned to the matter of picking up oxygen where it is abundant (lungs) and dispensing it readily where it is scarce (tissues). Closely related to these mechanisms are the reactions that take place when the abnormal hemoglobin, *hemoglobin S*, is deoxygenated. The consequences of the reactions that follow deoxygenation of *hemoglobin S* are responsible for the effects of the physiological traits, *sickle-cell trait* and *sickle-cell anemia*, the latter a serious debilitator of black Americans. Consult your biology teacher for ways in which the two of you can coordinate discussions to present the inorganic chemistry (page 96, student module*); the biochemistry (see *Molecules in Living Systems: A Biochemistry Module*, page 113); and the physiology, genetics, and evolution of this disease.**

I-45 REDOX ACTION

The conversion of food into energy is often spoken of as a process of burning the food. In a sense, this is an apt comparison. Ordinary burning is a rapid oxidation process that produces heat and light. The oxidation of food in the body also produces energy, but the processes obviously must differ considerably. Use of food must take place at body temperature and must be controlled to release energy only as the body requires it. The biological catalyst cytochrome c is involved in this process. Because it can be readily oxidized and reduced, it transfers electrons and makes the oxidation process possible.

In normal body metabolism, carbohydrates liberate energy at the rate of 4 kcal/g; proteins, also at 4 kcal/g; and fats, at the rate of 9 kcal/g* (see *Molecules In Living Systems: A Biochemistry Module*).

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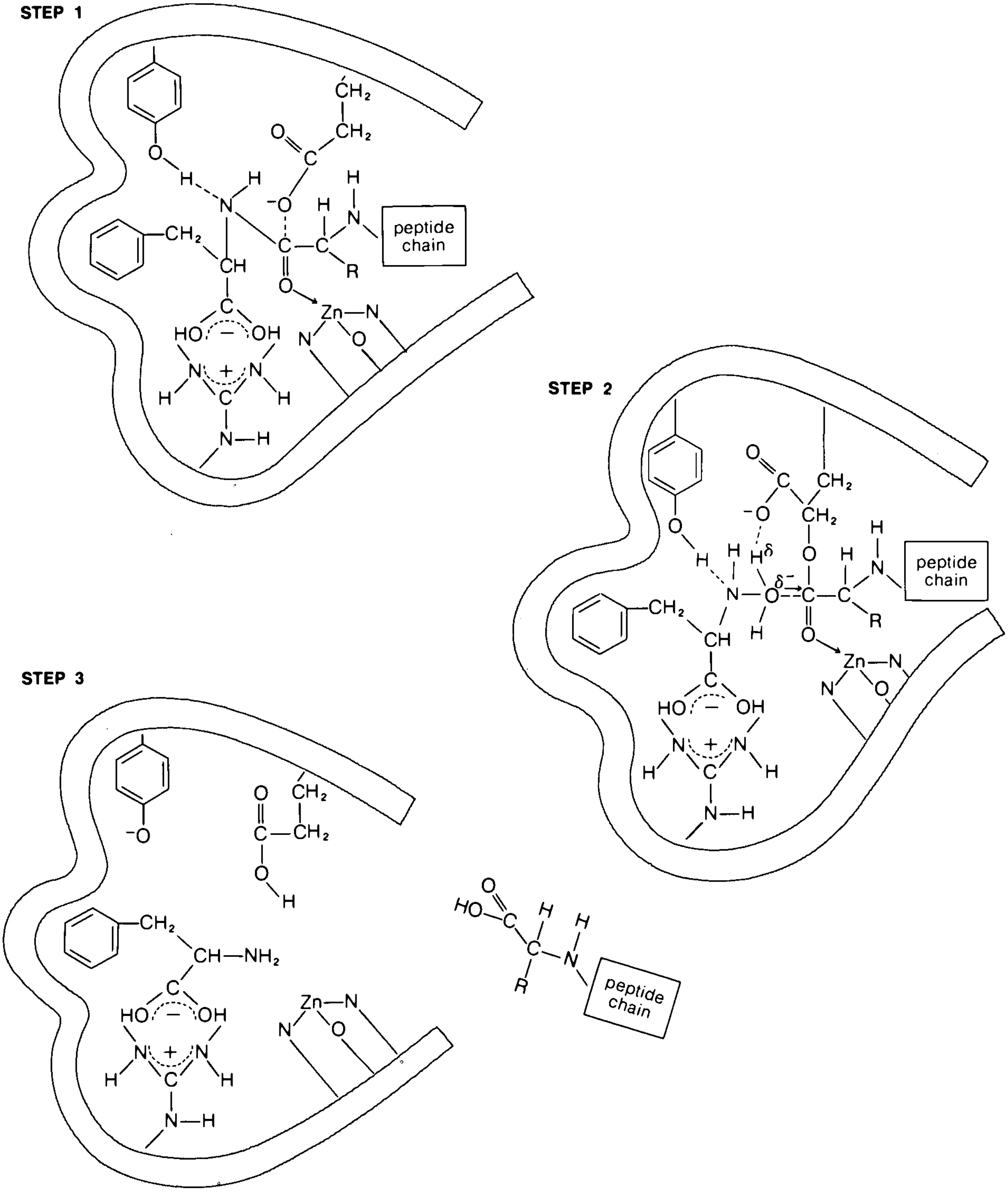
*See also J. E. Huheey, "Precious Metals—Life's Essentials," *REACTS-1973, Proceedings of the Regional Educators Annual Chemistry Teaching Symposium*. College Park, Md., Department of Chemistry, University of Maryland, and J. E. Huheey, *Inorganic Chemistry: Principles of Structure and Reactivity*, 2nd ed., (New York: Harper & Row, Publishers, 1978).

**See D. L. Martin and J. E. Huheey, *Journal of Chemical Education* Vol. 49 (1972), p. 177.

†The kcal, or kilocalorie, of the chemist is equivalent to the Calorie of the nutritionist and the weight-watcher.

MECHANISM OF ACTION OF CARBOXYPEPTIDASE A



ANSWERS TO PROBLEMS

(Student module page 96)

Fe—hemoglobin, myoglobin, cytochrome c, all used in respiration.
O_2—used in respiration by all animals.
Hg—poisonous element.
CO—coordinates preferentially to hemoglobin, causing asphyxia.
Zn—essential element in many enzymes, including carboxypeptidase A.
Pb—poisonous element.
NO—catalyzes the decomposition of O_3.
SO_2—pollutant from roasting or burning high-sulfur coal, and so forth; turns to sulfuric acid in the atmosphere.

EVALUATION ITEMS

These are evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. An inorganic air pollutant of much concern is

   A. CuS B. SO_2 C. CH_4 D. H_2O

2. The number of elements essential to all living systems is about

   A. 11 B. 87 C. 20 D. 42

3. A fertilizer that would be the least likely to include trace metals would be

   A. fish emulsion.
   B. manure.
   C. compost made from seaweed.
   D. ammonium nitrate.

4. A biological catalyst is

   A. an enzyme.
   B. an amino acid.
   C. a carbohydrate.
   D. the N_2 molecule.

5. The sole function of cytochrome c is

   A. to be a chemical meat cleaver.
   B. to be continuously oxidized and reduced.
   C. to carry oxygen through the blood.
   D. to store oxygen for transport.

A Summing Up

This summary provides the chance to discuss the intent of this module—to obtain an idea of what inorganic chemistry is and what an inorganic chemist does. A brief review of inorganic chemistry as presented in this module should be balanced by discussions in which students attempt to predict the future activities and value of inorganic chemistry.

ANSWERS TO PROBLEMS

(Student module page 99)

1. **a.** Acids: HCl, H_2SO_4, HNO_3, H_3PO_4; bases: NaOH, KOH, NH_3, etc.
   b. NaCl, CsCl, $CuSO_4$, etc.
   c. H_2, CH_4, NH_3, etc.
   d. BF_3, H^+, $AlCl_3$, etc.
   e. OH^-, Cl^-, NH_3, etc.
   f. $Cu(NH_3)_4^{2+}$, $Co(SCN)_4^{2-}$, etc.
   g. Fe, Ca, K, etc.
   h. Hg, Pb, etc.

2. **a.** $H_2SO_4 + 2KOH \longrightarrow K_2SO_4 + 2H_2O$
   b. $2Na + Cl_2 \longrightarrow 2NaCl$
   c. $HCO_3^- + OH^- \longrightarrow CO_3^{2-} + H_2O$
   d. $HCO_3^- + H_3O^+ \longrightarrow CO_2 + 2H_2O$
   e. $P_4 + 10Cl_2 \longrightarrow 4PCl_5$

3. **a.** Any of the examples discussed in *I-25* through *I-29*
   b. See section on *Bioinorganic Chemistry*
   c. Conduction of heat and electricity, malleability and ductility, and mechanical strength

4. $_{107}X$ might be isolated as KXO_4 (like $KMnO_4$). It will be a metal (fourth-row transition element) and will most closely resemble rhenium.

Appendix

Safety

SAFETY IN THE LABORATORY

Proper conduct in a chemistry laboratory is really an extension of safety procedures normally followed each day around your home and in the outside world. Exercising care in a laboratory demands the same caution you apply to driving a car, riding a motorbike or bicycle, or participating in a sport. Athletes consider safety measures a part of playing the game. For example, football players willingly spend a great deal of time putting on equipment such as helmets, hip pads, and shoulder pads to protect themselves from potential injury.

Chemists must also be properly dressed. To protect themselves in the laboratory, they commonly wear a lab apron or a coat and protective glasses. Throughout this course you will use similar items. Hopefully their use will become second nature to you, much as it becomes second nature for a baseball catcher to put on a chest protector and mask before stepping behind home plate.

As you read through a written experimental procedure, you will notice that specific hazards and precautions are called to your attention. Be prepared to discuss these hazards with your teacher and with your fellow students. Always read the entire experimental procedure thoroughly before starting any laboratory work.

A list of general laboratory safety procedures follows. It is not intended that you memorize these safety procedures but rather that you *use* them regularly when performing experiments. You may notice that this list is by no means complete. Your teacher may wish to add safety guidelines that are relevant to your specific classroom situation. It would be impossible to anticipate every hazardous situation that might arise in the chemistry laboratory. However, if you are familiar with these general laboratory safety procedures and if you use common sense, you will be able to handle potentially hazardous situations intelligently and safely. Treat all chemicals with respect, not fear.

GENERAL SAFETY GUIDELINES

1. Work in the laboratory only when the teacher is present or when you have been given permission to do so. In case of accident, notify your teacher immediately.
2. Before starting any laboratory exercise, be sure that the laboratory bench is clean.
3. Put on a laboratory coat or apron and protective glasses or goggles before beginning an experiment.
4. Tie back loose hair to prevent the possibility of its contacting any Bunsen burner flames.
5. Open sandals or bare feet are not permitted in the laboratory. The dangers of broken glass and corrosive liquid spills are always present in a laboratory.
6. Fire is a special hazard in the laboratory because many chemicals are flammable. Learn how to use the fire blanket, fire extinguisher, and shower (if your laboratory has one).
7. For minor skin burns, immediately immerse the burned area in cold water for several minutes. Then consult your teacher for further instructions on possible additional treatment.
8. In case of a chemical splash on your skin, immediately rinse the area with cold water for at least one minute. Consult your teacher for further action.
9. If any liquid material splashes into your eye, wash the eye immediately with water from an eyewash bottle or eyewash fountain.
10. Never look directly down into a test tube—view the contents of the tube from the side. (Why?)
11. Never smell a material by placing your nose directly at the mouth of the tube or flask. Instead, with your hand, "fan" some of the vapor from the container toward your nose. Inhale cautiously.
12. Never taste any material in the laboratory.
13. Never add water to concentrated acid solutions. The heat generated may cause spattering. Instead, as you stir, add the acid slowly to the water or dilute solution.
14. Read the label on a chemical bottle at least *twice* before removing a sample. H_2O_2 is not the same as H_2O.
15. Follow your teacher's instructions or laboratory procedure when disposing of used chemicals.

This symbol represents three of the common hazards in a chemistry laboratory—flame, fumes, and explosion. It will appear with certain experiments in this module to alert you to special precautions in addition to those discussed in this Appendix.

Suggested Readings

Note the references to books and articles that are listed as footnotes and in the text of this teacher's guide. You may also wish to consult the list of suggested readings published in the student module. For further reading into inorganic chemistry, you can refer to four inorganic textbooks at the college level:

Cotton, F. A., and Wilkinson, G. *Advanced Inorganic Chemistry*, 3rd ed. New York: John Wiley & Sons, 1972.

Demitras, G. C.; Russ, C. R.; Salmon, F.; Weber, J. H.; and Weiss, G. S. *Inorganic Chemistry*. Englewood Cliffs, N.J.: Prentice-Hall, 1972.

Huheey, J. E. *Inorganic Chemistry: Principles of Structure and Reactivity*, 2nd ed. New York: Harper & Row, 1978.

Purcell, K. F., and Kotz, J. C. *Inorganic Chemistry*. Philadelphia: W.B. Saunders Company 1977.

Two collections of articles from the *Journal of Chemical Education* are available:

Galloway, G., ed. *Collected Readings in Inorganic Chemistry*. Easton, Pa.: Chemical Education Publishing Company, 1972.

Watt, G. W., and Kieffer, W. F., eds. *Collected Readings in Inorganic Chemistry*. Easton, Pa.: Chemical Education Publishing Company, 1962.

In addition, see these articles in the University of Maryland REACTS series (Regional Educators Annual Chemistry Teaching Symposium, held each January at the University of Maryland, College Park): "New Compounds and New Ideas" (1970) and "Precious Metals—Life's Essentials" (1973).

Module Tests

Two module tests follow: one to test knowledge-centered objectives and the other to test skill-centered objectives. If you choose to use either or both of these module tests as they are presented here, duplicate copies for your students. Or, you may wish to select some of the questions from these tests that you feel apply to *your* introductory chemistry course and add additional questions of your own. Either way, make sure that the test you give reflects your emphasis on the chemistry that you and your students experienced in this module. The skill-centered tests will require that you set up several laboratory stations containing materials for your students to examine or work with. You may wish to add additional test items to round out the types of skills you and your students have worked on. (Answers to the test questions in this section are provided.) If you wish to use a standard-type answer sheet for this test, one is provided in the appendix of the teacher's guide for *Reactions and Reason: An Introductory Chemistry Module*.

ANSWERS FOR KNOWLEDGE-CENTERED MODULE TEST

1. C; **2.** A; **3.** B; **4.** D; **5.** D; **6.** D; **7.** B; **8.** A; **9.** D; **10.** C; **11.** C; **12.** D; **13a.** A; **13b.** B; **13c.** B; **14.** A; **15.** C; **16.** D; **17.** C; **18.** A; **19.** D; **20.** C; **21a.** B; **21b.** B; **22.** A; **23.** B; **24.** C; **25.** A

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DIVERSITY AND PERIODICITY
Knowledge-Centered Module Test

1. A base that is used in household cleaners is:

A. HCl B. $NaHSO_4$ C. NaOH D. $Ca(OH)_2$

2. The HCO_3^- ion can act either as an acid or a base. The reaction that best illustrates its action as a base is:

A. $HCO_3^- + H_2O \longrightarrow H_2CO_3 + OH^-$
B. $HCO_3^- + H_2O \longrightarrow CO_3^{2-} + H_3O^+$
C. $HCO_3^- + OH^- \longrightarrow CO_3^{2-} + H_2O$
D. $HCO_3^- + NH_3 \longrightarrow CO_3^{2-} + NH_4^+$

3. The formation of a coordination compound is represented by:

A. $Ca^{2+} + 2Cl^- \longrightarrow CaCl_2$
B. $CrCl_3 + 6\ NH_3 \longrightarrow Cr(NH_3)_6Cl_3$
C. $Cu^{2+} + SO_4^{2-} \longrightarrow CuSO_4$
D. $2Na + Cl_2 \longrightarrow 2NaCl$

4. A property of metals that is a result of their lattice structure is

A. their number of valence electrons.
B. their low ionization energies.
C. their ability to dissolve in acids.
D. their malleability.

5. Coordination compounds have uses

A. in qualitative analysis. C. in enzymes.
B. as water softeners. D. all of the above

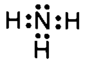
6. A student determined the ionization energies of four elements. Their ionization energies relative to hydrogen are as follows:

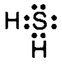
$H \longrightarrow H^+ + e^-$ I.E. = 1.00
$A \longrightarrow A^+ + e^-$ I.E. = 1.32
$B \longrightarrow B^+ + e^-$ I.E. = 1.62
$C \longrightarrow C^+ + e^-$ I.E. = 3.00
$D \longrightarrow D^+ + e^-$ I.E. = 0.32

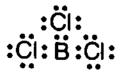
The element most likely to be a metal is:

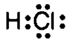
A. A B. B C. C D. D

7. An example of a Lewis acid is:

A.

$\mathrm{H{:}\underset{\displaystyle H}{\underset{\cdot\cdot}{\ddot{N}}}{:}H}$

C.

$\mathrm{H{:}\underset{\displaystyle H}{\underset{\cdot\cdot}{\ddot{S}}}{:}}$

B.

$\mathrm{\underset{:\ddot{C}l:\underset{\cdot\cdot}{B}:\ddot{C}l:}{:\ddot{C}l:}}$

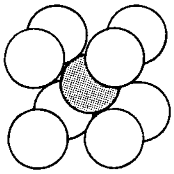
D.

$\mathrm{H{:}\underset{\cdot\cdot}{\ddot{C}l}{:}}$

8. Reduction is represented by:

A. $Mn^{3+} + e^- \longrightarrow Mn^{2+}$
B. $Na \longrightarrow Na^+ + e^-$
C. $H_3O^+ + OH^- \longrightarrow 2\ H_2O$
D. $K^+ + Cl^- \longrightarrow KCl$

9. The coordination number of the darkened ion in the following figure is:

A. 1 B. 4
C. 5 D. 8



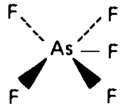
10. The family of elements requiring one electron to achieve a noble gas structure is called the

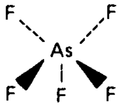
A. transition metals. C. halogens.
B. alkali metals. D. alkaline earth metals.

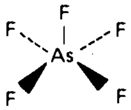
11. An ionic bond is

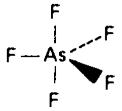
A. the unequal sharing of electrons between two atoms.
B. the equal sharing of electrons between two atoms.
C. the transfer of electrons from one atom to another.
D. the bond formed between two nonmetals.

12. If the compound formed between As and F has the formula AsF_5, the predicted geometry of the molecule is:

A.


B.


C.


D.


13. In a titration 365 grams of a household cleaner containing hydrochloric acid were neutralized with 100 cm^3 of a 1.0 M sodium hydroxide solution.

a. The balanced equation for this reaction is:

A. $NaOH + HCl \longrightarrow NaCl + H_2O$
B. $Na(OH)_2 + 2HCl \longrightarrow NaCl_2 + 2H_2O$
C. $Na_2OH + HCl \longrightarrow Na_2Cl + H_2O$
D. $NaOH + 2HCl \longrightarrow NaCl_2 + H_2O$

b. The number of moles of NaOH required was:

A. 0.01 B. 0.1 C. 1.0 D. 100

c. The percentage of hydrochloric acid in the cleaner was:

A. 0.1% B. 1% C. 8% D. 10%

14. $Co(NH_3)_6^{3+}$ is a complex that has a(an)

A. octahedral structure.
B. linear structure.
C. tetrahedral structure.
D. square planar structure.

15. Two properties of an element were found to be its ability to conduct heat well and its ductility. This element is most likely a

A. nonmetal. C. metal.
B. metalloid. D. noble gas.

16. Consider the following statements.

(a) Solid ionic compounds conduct electricity well.
(b) Solid ionic compounds are nonconductors.
(c) Liquid ionic compounds conduct electricity well.
(d) Liquid ionic compounds are nonconductors.
(e) Solutions of ionic compounds conduct electricity well.
(f) Solutions of ionic compounds are nonconductors.

Which of the above statements are true?

A. (a), (c), and (e) C. (a), (c), and (f)
B. (b), (d), and (e) D. (b), (c), and (e)

17. One hundred six elements are currently known. If element 107 is discovered, it will be similar to element 75. Element 114 will most likely be similar to:

A. At B. Ce C. Pb D. Pt

18. The octet rule is used to

A. predict a molecule's formula.
B. predict a molecule's conductivity.
C. estimate an atom's size.
D. predict an atom's coordination number.

19. An example of a biological function in which metal compounds are known to be involved is

A. oxygen transport in the blood.
B. electron transport reactions.
C. oxygen storage in muscles.
D. all of these

20. The two major divisions of the periodic table are the

A. nonmetals and the noble gases.
B. halogens and the nonmetals.
C. metals and the nonmetals.
D. metals and the metalloids.

21. In an experiment similar to the metal packing experiment, a student determined that he or she could stack 144 balls with a radius of 0.50 cm into a rectangular box 8 × 6 × 3 cm. (Note that the volume of a sphere is $1.33\pi r^3$.)

a. The volume occupied by the spheres is:

A. 48.1 cm^3 C. 101 cm^3
B. 75.4 cm^3 D. 144 cm^3

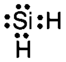
b. The calculated packing efficiency is:

A. 33.2% B. 52.3% C. 70.2% D. 100%

22. A physical property that is periodic in its nature is

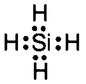
A. an element's ionization energy.
B. an element's radioactivity.
C. an element's color.
D. all of these

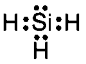
23. The correct electron-dot representation for the compound formed between Si and H is:

A.

$:\underset{H}{\overset{\cdot\cdot}{\underset{\cdot\cdot}{Si}}}:H$

C.

$H:\overset{\cdot\cdot}{\underset{\cdot\cdot}{Si}}:H$

B.

$H:\overset{H}{\underset{H}{Si}}:H$

D.

$H:\overset{\cdot\cdot}{\underset{H}{Si}}:H$

24. At room temperature CO_2 is a gas and H_2O is a liquid. This difference in physical state is a result of

A. the different acidic strengths of these molecules.
B. the different molecular sizes of these molecules.
C. the different molecular structures of these molecules.
D. the different molecular masses of these compounds.

25. Three drops of phenolphthalein indicator are added to an unknown solution and the solution turns red. The solution most likely contains:

A. KOH B. KF C. NaCl D. H_2SO_4

Skill-Centered Module Test

Using the skill-centered test items will require certain advance preparations on your part. The numerals in the following list indicate the items for which you will have to prepare special laboratory stations. Be sure to test each of the lab stations before allowing students to determine the answers to the skill-centered items. When students are ready to answer these questions, they should go to the numbered station and follow the directions that are given there and in the printed question item. When they finish with the materials at the station, instruct them to leave the materials in proper order for the next student.

2 Provide four samples of metals labeled A, B, C, and D. Be sure they are all of the approximate same size and *thickness*.

A is copper.	C is zinc or lead.
B is tin.	D is iron or nickel.

3 Using the plastic boxes, construct four packing models, making sure that only one is most efficient.

6 Fill a 50-cm^3 buret labeled A with 0.2 M HCl. Dilute 10 cm^3 of concentrated HCl to a final volume of 500 cm^3. Using a funnel, place extra solution in a stoppered bottle.

Supply a 125- or 250-cm^3 Erlenmeyer flask. Supply a dropper bottle of phenolphthalein. Fill buret B with 0.1 M NaOH. Dissolve 2 g NaOH in 500 cm^3 H_2O. Place extra in a stoppered bottle marked *base* using a funnel marked *for base use only*.

Titrate the sample beforehand so you know the amount needed.

7 Provide a solution of 0.1 M NaH_2PO_4. Place pH paper of appropriate range at the station with a color comparison chart. Be sure to measure the pH yourself.

8 Pieces of different metals marked with code letters or numbers in red crayon make good objects to weigh. Be sure to weigh them yourself.

9 Small rubber stoppers #0, 1, 2 make good objects for this determination. Mark them with code letters or numbers in red crayon. Provide a graduated cylinder (25- or 50-cm^3) and a beaker of water.

10 Provide three large test tubes labeled 1, 2 and 3. Place 15 cm^3 of H_2O in each. Add two drops of a food color to tube 1, four drops to tube 2, and six drops to tube 3.

Provide a large test tube labeled X. Place 15 cm^3 of H_2O in the tube and add five drops of food color to this tube.

12 Provide three test tubes labeled A, B, and C. Mark the 5 cm^3 level on each with a ring. Provide corks that will fit the test tubes. Provide three beakers labeled A, B, and C, containing the following:

A. 0.1 M NaCl	0.59 g/100 cm^3 H_2O
B. 0.1 M NaBr	1.05 g/100 cm^3 H_2O
C. 0.1 M NaI	1.50 g/100 cm^3 H_2O

Provide solution X (0.1 M NaI) in a beaker labeled X.

Provide chlorine water in a stoppered flask.

Provide hexane, another nonpolar solvent, in a stoppered flask.

Label this flask *solvent*.

Place a waste disposal jar for use in emptying the test tubes.

Chlorine water may be made by bubbling Cl_2 through water or by adding a small amount of acid (HCl) to a chlorine bleach (Clorox). Test this experiment before using, to be sure the solutions are of needed concentration.

ANSWERS FOR SKILL-CENTERED MODULE TEST

1. B; **2.** C; **3.** *; **4.** C; **5.** B; **6.** *; **7.** *; **8.** *; **9.** *; **10.** D; **11.** C; **12.** C

*Evaluate according to teacher standard

DIVERSITY AND PERIODICITY
Skill-Centered Module Test

Several questions in this section require you to make observations and perform chemical manipulations. The stations where you will do these operations will be indicated by your teacher. If the station you are going to is being used, continue with the test and go back later.

1. Consider the data below for eight elements:

Element	*A*	*B*	*C*	*D*	*E*	*F*	*G*	*H*
Ionization Potential (*eV*)	600	650	700	500	550	660	400	450
Boiling Point (°C)	1000	750	500	850	600	450	700	450
Atomic Radius (Å)	0.80	0.75	0.70	0.70	0.65	0.60	0.60	0.55

If you want to place these elements into periods, the smallest number of periods (periods should represent cycles of periodic properties) would be:

A. 2 B. 3 C. 4 D. 1

2. Go to station #2 and examine the four samples of metals provided. The most malleable (can be shaped most easily) sample is:

A. A B. B C. C D. D

3. Go to station #3 and examine the four crystal packing models provided. The model that represents the most efficient packing system is:

A. A B. B C. C D. D

4. Using the following data, calculate the efficiency of the packing system.

	(*l*)	(*w*)	(*h*)
Dimensions of cell	= 8 cm ×	5 cm ×	3 cm

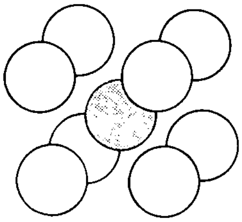
Diameter of ball = 1 cm
Number of balls in box= 100

The packing efficiency is:

A. 41.5% B. 92.5% C. 83.0% D. 33.3%

5. Examine the picture of a packing system below. The coordination number of the (central) colored atom is:

A. 6 B. 8 C. 10 D. 12



6. Go to station #6 and place 25 cm^3 of liquid cleaner from buret A into the 250-cm^3 Erlenmeyer flask. Add about 25 cm^3 of water. Add 2 drops of phenolphthalein indicator. Now titrate the sample with the base in buret B. Record the volume of base needed to neutralize the liquid cleaner next to #6 at the bottom of your answer sheet.

7. Go to station #7 and use the pH paper provided to measure the pH of the solution. Record your answer next to #7 at the bottom of your answer sheet.

8. Go to station #8 and select one of the objects provided and determine its mass. Record its identification number and its mass next to #8 at the bottom of your answer sheet.

9. Go to station #9 and determine the volume of one of the objects provided, using the water displacement method. Record its identification number and its volume next to #9 at the bottom of your answer sheet.

10. Go to station #10 and examine the three test tubes labeled 1, 2, and 3 in the test-tube rack. Now examine the test tube marked X. The color of X in comparison to the other three tubes is best described as:

A. same as tube #1
B. same as tube #2
C. between tubes #1 and 2
D. between tubes #2 and 3

11. A student synthesized several coordination compounds and made observations on their colors. These observations are summarized below.

Metal Ion	*Cl⁻*	*Ligand acetyl-acetone*	*NH₃*
Cu^{2+}	blue	black	blue
Zn^{2+}	white	black	white
Co^{2+}	red	black	red
Ni^{2+}	green	black	green

In order to complete the laboratory report, the student had to answer the following question. What factor(s) influence the colors of coordination compounds? The student should answer:

A. the ligand
B. the metal ion
C. both the ligand and the metal ion
D. neither the ligand nor the metal ion

12. Go to station #12 where you will find three test tubes labeled A, B, and C in the test-tube rack. The 5-cm^3 level is indicated by a ring. Fill each test tube with the liquid of the same letter. Now add about the same amount (5 cm^3) of chlorine water to each tube. Finally, add 2 cm^3 of solvent to each test tube and shake gently. Do not put your finger over the top of the tube, but use the cork provided. Now place 5 cm^3 of liquid X in tube X, add 5 cm^3 of chlorine water and 2 cm^3 of solvent, and shake gently, using the cork provided.

Solution X reacts like:

A. solution A
B. solution B
C. solution C
D. none of the above

EMPTY ALL TEST TUBES INTO THE WASTE DISPOSAL JAR BEFORE LEAVING.

Materials List

Quantities listed are for a class of 30 students working in pairs.

*Optional Items. These items depend on teacher choice. We have listed substitutions in the experiment discussion. Consult the specific experiment in the teacher's guide to determine use and quantities.

NONEXPENDABLE MATERIALS

Item	*Experiment*	*Amount*
Aluminum wire	12*	150 cm*
Aspirators	32	15
Balances, 0.01-g sensitivity	—	—
Balls, Styrofoam, 25-mm diameter	14, 19	2500–3000
Balls, Styrofoam, 50-mm diameter	14*, 19*	84*
Beakers, 50 cm^3	32	30
Beakers, 150 cm^3	32	15
Beakers, 250 cm^3	32	15
Beakers, 600 cm^3	32	15
Boxes, plastic, and false bottoms	14, 19	15
Büchner funnels, 91-mm diameter, and adapters	32	15
Bunsen burners	16, 20, 25, 32, 40	15
Burets, 50 cm^3	27	30
Buret clamps, double	27	15
Clamps, universal	32	15
Conductivity apparatus (dry cell, bulb, and leads)	12, 20	15 sets
Copper wire	12*	150 cm*
Corks, assortment	9, 35	x
Dropper bottles, 50 cm^3	38*	30*
Drying oven	32*	1*
Erlenmeyer flasks, 125 cm^3	32	15
Erlenmeyer flasks, 250 cm^3	27, 32	45
Filter flasks, 250 cm^3	32	15
Funnels, 75-mm diameter, and supports	27	15
Glass squares, 8 × 8 cm, clear (or clear plastic)	18	30
Graduated cylinders, 10 cm^3	9, 27, 35, 36	15
Graduated cylinders, 50 cm^3	32	15
Graphite	12*	5 g*
Hot plate	32*	1–10*
Iron wire	12*	150 cm*
Lead wire	12*	150 cm*
Magnets, disc, 1 cm	18, 23	100
Magnets, ring	18	10
Medicine droppers	38, 40	30
Model kits, molecular	23*	15*
Molybdenum wire	12*	150 cm*
Mortar and pestle sets	20*	15*
Overhead projector and screen	18*	1*
Petri dishes, Pyrex	16	30
Platinum wire	12*	30 cm*
Ring stands and rings	16, 32	15
Rubber stoppers, #6, 2-hole, to fit 250-cm^3 flasks (regular-size mouth)	32	15

Item	*Experiment*	*Amount*
Rubber stopper assortment	1	1
Rubber tubing, heavy wall, 9 mm (3/8″) I.D. (15 lengths)	32	900 cm
Scissors	4	15
Separatory funnels, 250 cm^3	38	3
Test tubes, 18 × 150 mm	9, 20, 35, 36, 38, 40	240
Test-tube clamps	40	15
Test-tube racks	35, 36, 38	15
Thermometers, −10°C to 110°C	32	15
Tungsten wire	12*	150 cm*
Vials, 30 cm^3	32	15–120
Watch glass, 90-mm diameter	32	30
Wire gauze, asbestos centers	16, 32	15
Zinc wire	12*	150 cm*

EXPENDABLE MATERIALS

Item	*Experiment*	*Amount*
Acetic acid, glacial	32, 40	150 cm^3
Acetone	32	800 cm^3
Acetylacetone, $CH_3COCH_2COCH_3$ (2,4-pentanedione)	32	600 cm^3
Aluminum nitrate, $Al(NO_3)_3 \cdot 9H_2O$	35	6 g
Aluminum sulfate, $Al_2(SO_4) \cdot 18H_2O$	32	90 g
Ammonia, conc., NH_3 (ammonium hydroxide, conc. NH_4OH)	32, 40	1000 cm^3
Ammonium aluminum sulfate, $NH_4Al(SO_4)_2 \cdot 12H_2O$	16	60 g
Ammonium chloride, NH_4Cl	32	150 g
Balloons, round	23	150
Bromine water	9	25 cm^3
Calcium nitrate, $Ca(NO_3)_2 \cdot 4H_2O$	35, 36	11 g
Charcoal, powdered	32	15 g
Chlorine water (or commercial chlorine bleach)	9	100 cm^3
Chromium chloride, hexahydrate, $CrCl_3 \cdot 6H_2O$	32	45 g
Cobalt carbonate, $CoCO_3$	32	45 g
Cobalt chloride, hexahydrate, $CoCl_2 \cdot 6H_2O$	32	500 g
Cobalt nitrate, hexahydrate, $Co(NO_3)_2 \cdot 6H_2O$	20*, 32, 35, 38	170 g
Copper(II) sulfate, pentahydrate, $CuSO_4 \cdot 5H_2O$	16, 20*, 32	600 g
Dithizone (di-phenol-thiocarbazone)	40	0.1 g
Drano, Vanish, Saniflush, etc.	27	1 can each
Ethanol (ethyl alcohol)	32	2500 cm^3
Ethylenediamine, $H_2NCH_2CH_2NH_2$	32	300 cm^3
Filter paper, medium weight, 9-cm diameter (for Büchner funnels)	32	200 sheets
Filter paper, 12.5-cm diameter	40	100
Glass tubing, 6-mm diameter	32	5 m
Graph paper, linear	4	30 sheets
Hydrobromic acid, HBr conc.	38	250 cm^3
Hydrochloric acid, HCl conc.	27, 32, 38	850 cm^3
Hydrogen peroxide, H_2O_2, 30 percent	32	1000 cm^3
Iodine, I_2	9, 12*	5 g
Iodine solution (I_2 in KI reagent)	9	50 cm^3

EXPENDABLE MATERIALS (cont'd)

Item	*Experiment*	*Amount*
Iron(III) chloride, hexahydrate, $FeCl_3 \cdot 6H_2O$	32	45 g
Iron(III) nitrate, $Fe(NO_3)_3$	35, 36, 38	10 g
Lead nitrate	40	5 g
Lithium, metal	12*	5 g*
Litmus, red and blue	32	15 vials each
Magnesium nitrate, dihydrate, $Mg(NO_3)_2 \cdot 2H_2O$	35, 36	12 g
Methanol (methyl alcohol)	25	10 cm^3
Methyl isobutyl ketone (4-methyl-2-pentanone or MIBK)	38	250 cm^3
Methylene chloride	14	250 cm^3
Nitrilotriacetic acid, sodium salt, $N(CH_3COONa)_3$ (NTA)	36	8 g
Phenolphthalein, 1 percent	27	10–50 cm^3
Potassium aluminum sulfate, $KAl(SO_4)_2 \cdot 12H_2O$	16, 20*	100 g
Potassium bromide, KBr	9, 38	100 g
Potassium chloride, KCl	9	1 g
Potassium dichromate, $K_2Cr_2O_7$	20*	10 g*
Potassium iodide, KI	9, 20*	2 g
Potassium nitrate, KNO_3	20*	10 g*
Potassium permanganate, $KMnO_4$	16	15 g
Soap (castile)	35, 36	6 g
Sodium, metal	12*	10 g*
Sodium carbonate, Na_2CO_3	36	3 g
Sodium chloride	20*	10 g*
Sodium hydroxide, NaOH	27	120 g
Sodium nitrate, $NaNO_3$	20*	10 g*
Sodium nitrite, $NaNO_2$	32	450 g
Sodium tetraborate, $Na_2B_4O_7$	36	6 g
Sodium thiocyanate, NaSCN	38	100 g
Sodium triphosphate, $Na_5P_3O_{10}$	36	10 g
Stopcock grease	27	1 tube
Sulfur, lump	12*	50 g*
Tape, clear (or glue)	4	1 roll
Toluene	20*, 32*	900 cm^3*
Trichlorotrifluoroethane (TTE, Freon 113, or Du Pont TF Solvent)	9, 40	300 cm^3
Urea, H_2NCONH_2	32	300 g
Vinegar	27	500 cm^3

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Table of International Relative Atomic Masses*

Element	Symbol	Atomic Number	Atomic Mass
Actinium	Ac	89	227.0
Aluminum	Al	13	27.0
Americium	Am	95	(243)**
Antimony	Sb	51	121.8
Argon	Ar	18	39.9
Arsenic	As	33	74.9
Astatine	At	85	(210)
Barium	Ba	56	137.3
Berkelium	Bk	97	(247)
Beryllium	Be	4	9.01
Bismuth	Bi	83	209.0
Boron	B	5	10.8
Bromine	Br	35	79.9
Cadmium	Cd	48	112.4
Calcium	Ca	20	40.1
Californium	Cf	98	(251)
Carbon	C	6	12.0
Cerium	Ce	58	140.1
Cesium	Cs	55	132.9
Chlorine	Cl	17	35.5
Chromium	Cr	24	52.0
Cobalt	Co	27	58.9
Copper	Cu	29	63.5
Curium	Cm	96	(247)
Dysprosium	Dy	66	162.5
Einsteinium	Es	99	(254)
Erbium	Er	68	167.3
Europium	Eu	63	152.0
Fermium	Fm	100	(257)
Fluorine	F	9	19.0
Francium	Fr	87	(223)
Gadolinium	Gd	64	157.3
Gallium	Ga	31	69.7
Germanium	Ge	32	72.6
Gold	Au	79	197.0
Hafnium	Hf	72	178.5
Helium	He	2	4.00
Holmium	Ho	67	164.9
Hydrogen	H	1	1.008
Indium	In	49	114.8
Iodine	I	53	126.9
Iridium	Ir	77	192.2
Iron	Fe	26	55.8
Krypton	Kr	36	83.8
Lanthanum	La	57	138.9
Lawrencium	Lr	103	(260)
Lead	Pb	82	207.2
Lithium	Li	3	6.94
Lutetium	Lu	71	175.0
Magnesium	Mg	12	24.3
Manganese	Mn	25	54.9
Mendelevium	Md	101	(258)
Mercury	Hg	80	200.6
Molybdenum	Mo	42	95.9
Neodymium	Nd	60	144.2
Neon	Ne	10	20.2
Neptunium	Np	93	237.0
Nickel	Ni	28	58.7
Niobium	Nb	41	92.9
Nitrogen	N	7	14.0
Nobelium	No	102	(259)
Osmium	Os	76	190.2
Oxygen	O	8	16.0
Palladium	Pd	46	106.4
Phosphorus	P	15	31.0
Platinum	Pt	78	195.1
Plutonium	Pu	94	(244)
Polonium	Po	84	(209)
Potassium	K	19	39.1
Praseodymium	Pr	59	140.9
Promethium	Pm	61	(145)
Protactinium	Pa	91	231.0
Radium	Ra	88	226.0
Radon	Rn	86	(222)
Rhenium	Re	75	186.2
Rhodium	Rh	45	102.9
Rubidium	Rb	37	85.5
Ruthenium	Ru	44	101.1
Samarium	Sm	62	150.4
Scandium	Sc	21	45.0
Selenium	Se	34	79.0
Silicon	Si	14	28.1
Silver	Ag	47	107.9
Sodium	Na	11	23.0
Strontium	Sr	38	87.6
Sulfur	S	16	32.1
Tantalum	Ta	73	180.9
Technetium	Tc	43	(97)
Tellurium	Te	52	127.6
Terbium	Tb	65	158.9
Thallium	Tl	81	204.4
Thorium	Th	90	232.0
Thulium	Tm	69	168.9
Tin	Sn	50	118.7
Titanium	Ti	22	47.9
Tungsten	W	74	183.8
Uranium	U	92	238.0
Vanadium	V	23	50.9
Xenon	Xe	54	131.3
Ytterbium	Yb	70	173.0
Yttrium	Y	39	88.9
Zinc	Zn	30	65.4
Zirconium	Zr	40	91.2

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*Based on International Union of Pure and Applied Chemistry (IUPAC) values (1975).

**Numbers in parentheses give the mass numbers of the most stable isotopes.

PERIODIC TABLE OF THE ELEMENTS

1.008
H
Hydrogen
1

200.6 — atomic mass
Hg — symbol
Mercury — name
80 — atomicnumber

IA	IIA	IIIB	IVB	VB	VIB	VIIB	← VIIIB →			IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
																	4.00 **He** Helium 2
6.94 **Li** Lithium 3	9.01 **Be** Beryllium 4											10.8 **B** Boron 5	12.0 **C** Carbon 6	14.0 **N** Nitrogen 7	16.0 **O** Oxygen 8	19.0 **F** Fluorine 9	20.2 **Ne** Neon 10
23.0 **Na** Sodium 11	24.3 **Mg** Magnesium 12											27.0 **Al** Aluminum 13	28.1 **Si** Silicon 14	31.0 **P** Phosphorus 15	32.1 **S** Sulfur 16	35.5 **Cl** Chlorine 17	39.9 **Ar** Argon 18
39.1 **K** Potassium 19	40.1 **Ca** Calcium 20	45.0 **Sc** Scandium 21	47.9 **Ti** Titanium 22	50.9 **V** Vanadium 23	52.0 **Cr** Chromium 24	54.9 **Mn** Manganese 25	55.8 **Fe** Iron 26	58.9 **Co** Cobalt 27	58.7 **Ni** Nickel 28	63.5 **Cu** Copper 29	65.4 **Zn** Zinc 30	69.7 **Ga** Gallium 31	72.6 **Ge** Germanium 32	74.9 **As** Arsenic 33	79.0 **Se** Selenium 34	79.9 **Br** Bromine 35	83.8 **Kr** Krypton 36
85.5 **Rb** Rubidium 37	87.6 **Sr** Strontium 38	88.9 **Y** Yttrium 39	91.2 **Zr** Zirconium 40	92.9 **Nb** Niobium 41	95.9 **Mo** Molybd'm 42	(97) **Tc** Technetium 43	101.1 **Ru** Ruthenium 44	102.9 **Rh** Rhodium 45	106.4 **Pd** Palladium 46	107.9 **Ag** Silver 47	112.4 **Cd** Cadmium 48	114.8 **In** Indium 49	118.7 **Sn** Tin 50	121.8 **Sb** Antimony 51	127.6 **Te** Tellurium 52	126.9 **I** Iodine 53	131.3 **Xe** Xenon 54
132.9 **Cs** Cesium 55	137.3 **Ba** Barium 56	138.9 **La*** Lanthanum 57	178.5 **Hf** Hafnium 72	180.9 **Ta** Tantalum 73	183.8 **W** Tungsten 74	186.2 **Re** Rhenium 75	190.2 **Os** Osmium 76	192.2 **Ir** Iridium 77	195.1 **Pt** Platinum 78	197.0 **Au** Gold 79	200.6 **Hg** Mercury 80	204.4 **Tl** Thallium 81	207.2 **Pb** Lead 82	209.0 **Bi** Bismuth 83	(209) **Po** Polonium 84	(210) **At** Astatine 85	(222) **Rn** Radon 86
(223) **Fr**† Francium 87	226.0 **Ra** Radium 88	(227) **Ac**** Actinium 89	‡ 104	‡ 105	‡ 106												

*	140.1 **Ce** Cerium 58	140.9 **Pr** Praseody'm 59	144.2 **Nd** Neodymium 60	(145) **Pm** Promethium 61	150.4 **Sm** Samarium 62	152.0 **Eu** Europium 63	157.3 **Gd** Gadolinium 64	158.9 **Tb** Terbium 65	162.5 **Dy** Dysprosium 66	164.9 **Ho** Holmium 67	167.3 **Er** Erbium 68	166.9 **Tm** Thulium 69	173.0 **Yb** Ytterbium 70	175.0 **Lu** Lutetium 71
**	232.0 **Th** Thorium 90	231.0 **Pa** Protactinium 91	238.0 **U** Uranium 92	237.0 **Np** Neptunium 93	(242) **Pu** Plutonium 94	(243) **Am** Americium 95	(245) **Cm** Curium 96	(245) **Bk** Berkelium 97	(251) **Cf** Californium 98	(254) **Es** Einsteinium 99	(254) **Fm** Fermium 100	(256) **Md** Mendel'm 101	(254) **No** Nobelium 102	(257) **Lr** Lawrencium 103

†The most stable known isotopes are shown in parentheses.

‡The discovery of elements 104, 105, and 106 has been claimed by both American and Soviet scientists. The Americans have suggested the name *rutherfordium* and *hahnium* for 104 and 105; the Soviets have suggested the names *kurchatovium* and *nielsbohrium* for these same elements. No name has yet been proposed for element 106.

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