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UNIT 1

The Diversity of Matter and Chemical Bonding

Teaching Unit 1: The Diversity of Matter and Chemical Bonding

Chapter 1 *Chemical Bonding*

Answers to Questions for Comprehension, notes on Figures and Web Links appear throughout the chapter.

Chapter 2 *Diversity of Matter*

Answers to Questions for Comprehension, notes on Figures and Web Links appear throughout the chapter.

Teaching Unit 1: The Diversity of Matter and Chemical Bonding

(20% of the course time; approximately 25 hours)

Student Textbook pages 2–89

General Outcomes

- describe the role of modelling, evidence, and theory used in explaining and understanding the structure, chemical bonding, and properties of ionic compounds
- describe the role of modelling, evidence, and theory used in explaining and understanding the structure, chemical bonding, and properties of molecular substances

Contents

Chapter 1 Chemical Bonding

Chapter 2 Diversity of Matter

Content Summary

Chemistry is the study of the physical and chemical properties of substances. These properties are determined by the chemical bonds between atoms in molecules, and by the forces of attraction between molecules. This unit introduces students to chemical bonding and explains how different bond types account for the diversity of properties we observe in matter.

The structure of matter is illustrated from the atomic level to the aggregate level. The electron structure of atoms dictates how they combine chemically to form bond types with a range of characteristics. The chemical formula of a compound can be predicted knowing the number of valence electrons present in the combining atoms, while the type of bond formed can be predicted using periodic properties. The three-dimensional structure of ionic crystals, and the strength of the ionic bond, explains the physical properties of ionic compounds. Explaining the more varied properties found in covalent compounds requires an understanding of molecular shapes, covalent bonding, and intermolecular forces.

Chapter 1 introduces students to ionic and covalent bonding. Students will use the octet rule and electron dot diagrams to predict the type of bonding between a pair of atoms. Using the charges on ions, the chemical formula of an ionic compound can be predicted. The chemical formulas of simple molecules and polyatomic ions can be predicted by writing Lewis structures. The delocalized electron model is used to explain metallic bonding and, in Chapter 2, this model is used to explain the properties of metals. The concept of electronegativity is introduced and the periodic trends of this property explained. The electronegativity difference between bonded atoms is used to help explain the bonding continuum from mostly ionic to purely covalent bonds.

In Chapter 2, the packing of ions into a crystal lattice is described. The VSEPR model is used to determine the three-dimensional shape of a covalent molecule, and this is combined with the concept of polar bonding to predict whether a molecule has a permanent dipole. Network solids are described as an array of covalently bonded atoms. Students are introduced to dipole-dipole attractions, hydrogen bonding, and London (dispersion) forces, and they learn how these intermolecular forces affect the physical properties of substances. The special role of hydrogen bonding in biological molecules and in water, hydrogen fluoride, and ammonia is illustrated. The differing electrical conductivity of ionic, covalent, and metallic substances is explained in terms of the flow of charged particles. Finally, students experiment with a variety of solids and use their knowledge to classify each substance as non-polar covalent, polar covalent, ionic, network, or metallic.

Curriculum Fit

Background: This first unit in *Chemistry 20* builds on *Science 9*, Unit 5, Chemical Properties of Change, and on *Science 10*, Unit C, Energy and Matter in Chemical Change. Students' understanding of the structures of atoms and molecules, the periodic properties of the elements, and chemical nomenclature are foundations upon which *Chemistry 20* builds. Students must also be familiar with the WHMIS unit, and a review of safe practices in the laboratory is recommended before activities begin.

The concepts introduced in Unit 1 are relevant to several of the units that follow in *Chemistry 20-30*. An understanding of intermolecular forces allows students to predict which substances will be gases at room temperature and pressure. Students will use intermolecular forces to predict which solute and solvent combinations are likely to form solutions. Both intramolecular and intermolecular forces are at work when an acid or a base dissolves in water, and the concept of polar bonding is used to determine which acids will be strong and which will be weak.

Core Concepts

Concept	Outcome	Text Reference
A chemical bond is the force of electrical attraction that holds two atoms together in a molecule or formula unit. The electrical attraction is between the nuclei of the bonded atoms and a pair of electrons shared between the nuclei.	20–A1.3k	Unit 1 Prep, pp. 8-9 Section 1.1, pp. 18-20; 23-29
Valence electrons are those electrons that are involved in the formation of chemical bonds. For the main group elements, the valence electrons are in the outer level.	20–A1.3k; 20–A1.4k; 20–A1.2s	Unit 1 Prep, p. 8 Section 1.1, p. 19
Covalent bonding is a type of bonding in which the electron pair is shared equally between the bonded nuclei. In a polar covalent bond the electron pair is shared unequally, so the bond is polar. Polar bonds have a partially charged negative pole and a partially charged positive pole.	20–A2.3k	Section 1.1, p. 23 Section 1.2, pp. 39-42
In ionic bonding, the bonding electron pair is shared so unequally that it is approximated as an electron transfer. As a result of the transfer of one or more electrons, a pair of ions is formed. The ionic bond is the electrical attraction between oppositely charged ions.	20–A1.3k; 20–A1.4k	Unit 1 Prep, p. 9 Section 1.1, pp. 19-20
The formula of an ionic compound is the simplest whole-number ratio of ions that results in a net charge of zero.	20–A1.2k	Section 1.1, p. 21
Electronegativity is a measure of the relative ability of an atom in a bond to attract the shared electrons. The difference in electronegativities between the atoms in a bond determines the bond type.	20–A1.3k; 20–A1.2sts; 20–A2.10k	Section 1.2, pp. 36-42
Electron dot diagrams are models to help visualize how valence electrons on atoms form bonds. The ionic bond is shown as an electric attraction between oppositely charged ions formed by the transfer of one or more electrons. Covalent bonds are shown as the sharing of one or more pairs of electrons.	20–A1.4k; 20–A1.5k; 20–A1.2s; 20–A2.3k; 20–A2.4k	Unit 1 Prep, pp. 8-9 Section 1.1, pp. 18-31
Lewis structures show how the valence electrons are shared among the atoms in a molecule.	20–A2.4k;	Section 1.1, pp. 23-29
A structural formula shows the number of atoms in a molecule, their position, and the bonds between them.	20–A2.4k	Section 1.1, p. 31
The formula of a molecular compound refers to the number of atoms of each element in the substance.	20–A2.2k	Unit 1 Prep, p. 11
The Valence-Shell Electron-Pair Repulsion (VSEPR) Model is used to predict the three-dimensional shape of a molecule by considering the structure that minimizes electron-pair repulsions.	20–A2.5k; 20–A2.6k	Section 2.1, pp. 52-56
A polar molecule has a permanent dipole moment as a result of its polar bonds and shape.	20–A2.9k	Section 2.1, pp. 57-59
Intramolecular forces are chemical bonds: ionic, covalent, or metallic bonds.	20–A1.3k	Section 2.2, p. 63
Ionic compounds are composed of ions in a three-dimensional lattice structure. This arrangement explains some of the properties of ionic substances.	20–A1.6k	Section 2.1, pp. 48-50

Core Concepts

Concept	Outcome	Text Reference
Observing the electrical conductivity of a substance as a solid and in aqueous solution can distinguish between an ionic, covalent, or metallic substance.	20–A2.8k	Section 2.3, pp. 76-77
Intermolecular forces are forces between molecules, atoms, or ions. These forces are generally classified as dipole-dipole, hydrogen bonding, or London (dispersion) forces.	20–A2.7k	Section 2.2, pp. 63-68
The strength of the intermolecular forces for a molecular substance is related to the physical properties of the substance.	20–A2.8k; 20–A2.1sts	Section 2.3, pp. 71-74

Beyond the Core Concepts

Concept	Outcome	Text Reference
Metallic bonds result from the electrical attraction between positively charged metal ions and delocalized valence electrons.		Section 1.1, p. 32 Section 2.3, p. 75
Network solids contain atoms joined by covalent bonds that extend throughout the crystal.		Section 2.1, pp. 60-61
Chemical compounds are important in many everyday processes and products.	20–A1.1sts	Unit Prep, pp. 14; Section 1.1, pp. 16-17 Section 1.1, p. 22 Section 2.2, pp. 69-70 Section 2.3, p. 79
New materials result from research on chemical bonding and the development of new technologies	20–A2.2sts; 20–A1.3sts	Section 2.3, pp. 84-85

Related Skills

Skill	Outcome	Text Reference
Given the name or formula of: • a binary ionic compound; • a binary covalent compound; or • an ionic compound containing polyatomic ions students should be able to write the formula or name of the compound.	20–A1.1k; 20–A2.1k	Unit Prep, pp. 10-13
Students must be able to describe procedures for safe handling, storage, and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information.	20–A1.1s	Safety in Your Chemistry Laboratory and Classroom, p. xiv
Students should be able to use the periodic table to make predictions about the type of bonding and the nomenclature of binary compounds.	20–A1.2s	Section 1.1, pp. 19-21
Students will build models of different crystal structures and compare them with models built by other groups.	20–A1.2s 20–A1.4s 20–A2.4s	Section 2.1, pp. 51

Related Skills

Skill	Outcome	Text Reference
Students will model some three-dimensional features of simple molecules, including polar bonds.	20–A2.6k 20–A2.7k	Section 1.1, p. 30 Section 2.1, pp. 56-57 Section 2.2, p. 64
Students will perform experiments to compare the properties of molecular compounds and to identify an unknown as either non-polar covalent, polar covalent, ionic, network or metallic	20–A2.2s 20–A1.3s	Section 2.2, pp. 66-67 Section 2.3, p. 78
Students will predict the physical properties of molecular substances based on the strength of the intermolecular forces present.	20–A2.1s	Section 2.3, pp. 71-74
Students will analyze trends in the physical properties of a related series of molecular substances and make a hypothesis about the intermolecular forces causing the trend	20–A2.3s	Section 2.3, p. 80

Activities and Target Skills

Activity	Target Skills
Chapter 1: Chemical Bonding	
Launch Lab: Chemistry Recall, p. 15	■ Naming compounds and identifying types of bonds
Thought Lab 1.1: Lewis Structures, p. 27	■ Drawing Lewis structures
Investigation 1.A: Modelling Molecules, p. 30	■ Drawing electron dot diagrams and building models ■ Analyzing and evaluating models constructed by others
Chapter 2: Diversity of Matter	
Launch Lab: Crystalline Columns, p. 47	■ Observing the growth of crystals ■ Explaining how stalagmites and stalactites could form
Investigation 2.A: Building Ionic Crystals, p. 51	■ Building models of common ionic lattices ■ Working collaboratively ■ Critically analyzing models of ionic compounds built by others
Investigation 2.B: Soap Bubble Molecules, pp. 56-57	■ Building models of molecular substances ■ Evaluating and analyzing models constructed by others
Investigation 2.C: Dipole Balloons, p. 64	■ Analyzing models of molecular substances
Investigation 2.D: Investigating the Properties of Water, pp. 66-67	■ Explaining the properties of molecular substances based on attractive forces
Investigation 2.E: Properties of Substances, p. 78	■ Comparing the properties of compounds ■ Analyzing experimental data to determine the properties of compounds ■ Determining the relative melting points of substances

Conceptual Challenges

Section 1.2

- Calculating the difference in the electronegativity (ΔEN) between two bonding atoms tells us whether the bond formed between these atoms will be ionic, polar covalent, or non-polar covalent. The ΔEN refers only to an individual bond and not to the compound itself. That is, the ΔEN is an indicator of bond polarity, not molecular polarity. The polarity of a molecule results from an analysis of all the bonds in a molecule. Students with an understanding of vector addition will appreciate that the molecular dipole represents the vector sum of all the bond dipoles in the molecule.

Section 2.1

- Ionic compounds do not exist as molecules but as anions and cations together in a large, three-dimensional crystal lattice. It is important to emphasize that not only are ionic bonds strong but they extend from one ion to another throughout the crystal. Ball and stick models of ionic crystals can give students an incorrect idea about the ionic bond. It is natural to associate the joining of ions in such models as showing a directional bond, but the ionic bond is not directional. For this reason, carefully monitor the models of ionic crystals that students build during Investigation 2.A. Ideally, the toothpicks used to join the polystyrene balls should not be visible when the model is completed.

Using the Unit 1 Preparation Feature

The unit opener has been designed to establish a social, technical, and environmental context for the science in the unit. Use the unit opener to introduce the general unit topics within that context, and ask the Focussing Questions to guide students' thinking.

The Unit Preparation feature has been included in order to ensure that students are familiar with the science from previous courses that relates specifically to the material they are about to study. Encourage students to take the Unit Prequiz (found at **www.albertachemistry.ca**, Online Learning Centre, Student Edition) to gauge their recall, noting that if they are familiar with the background science, their experience with this unit will be much easier.

The Unit 1 Preparation feature is one of the most extensive in *Inquiry into Chemistry* and is designed to refresh the memories of returning students. It covers 10 pages, and topics reviewed include: atoms, elements, and isotopes; the periodic table; electron energy levels; electron dot diagrams; forming bonds; names and formulas for binary compounds; and ionic compounds with polyatomic ions. The feature also includes exercises to help students further test their recall.

Answers to Questions for Comprehension

Student Textbook page 7

Q1.

Chemical notation	Element	Atomic number (Z)	Mass number (A)	Number of protons	Number of neutrons
$^{11}_{5}\text{B}$	boron	5	11	5	6
$^{208}_{82}\text{Pb}$	lead	82	208	82	126
$^{184}_{74}\text{W}$	tungsten	74	184	74	110
$^{4}_{2}\text{He}$	helium	2	4	2	2
$^{239}_{94}\text{Pu}$	plutonium	94	239	94	145
$^{56}_{26}\text{Fe}$	iron	26	56	26	30
$^{209}_{83}\text{Bi}$	bismuth	83	209	83	126
$^{107}_{47}\text{Ag}$	silver	47	107	47	60*
$^{20}_{10}\text{Ne}$	neon	10	20	10	10
$^{108}_{47}\text{Ag}$	silver	47	108	47	61
$^{75}_{33}\text{As}$	arsenic	33	75	33	42
$^{80}_{35}\text{Br}$	bromine	35	80	35	45
$^{197}_{79}\text{Au}$	gold	79	197*	79	118
$^{119}_{50}\text{Sn}$	tin	50	119	50	69

*Please use these revised numbers.

Q2. (a) helium, neon, argon, krypton, xenon, radon

(b) beryllium, magnesium, calcium, strontium, barium, radium

(c) fluorine, chlorine, bromine, iodine, astatine

(d) lithium, sodium, potassium, rubidium, cesium, francium

Answers to Questions for Comprehension

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Q3. (a) aluminium(III) oxide

(b) mercury(II) iodide

(c) sodium phosphide

(d) calcium bromide

Q4. (a) ZnO

(b) FeS

(c) MgI_2

(d) $CoCl_3$

Q5. (a) sulfur hexafluoride

(b) dinitrogen pentoxide

(c) phosphorus pentachloride

(d) carbon tetraflouride

Q6. (a) H_2O

(b) SO_3

(c) N_2O_4

(d) N_2O

Q7. (a) potassium phosphate

(b) ammonium chloride

(c) lithium perchlorate

(d) sodium hydrogen carbonate

Q8. (a) $KClO$

(b) $MgOOCCOO$

(c) $NaCN$

(d) $(NH_4)_2SO_4$

UNIT 1: COURSE MATERIALS

Chapter, Section	Item Description	Suggested Quantity (assume 40 in class)	Text Activity
Chapters 1, 2	safety goggles	40	Chapter 2 Launch Lab; Investigations: 2.B, 2.E
Chapters 1, 2	nonlatex disposable gloves	40 pairs × 4 investigations	Chapter 2 Launch Lab; Investigations: 2.B, 2.E
Chapters 1, 2	aprons	40	Chapter 2 Launch Lab; Investigations: 2.B, 2.E
Chapter 1, Section 1.1	pencil paper cut in 2 cm × 2 cm squares	1 per student 1 sheet per student	Launch Lab: Chemistry Recall, p. 15
Chapter 1, Section 1.1	molecular modelling kit	1 per group	Investigation 1.A: Modelling Molecules, p. 30
Chapter 2, Section 2.1	balance 10 mL graduated cylinder 100 mL Erlenmeyer flask water sodium acetate trihydrate crystals hot plate water bottle burette and stand glass plate forceps	1 per pair 1 per pair 1 per pair 5 mL per pair 55 g per pair 1 per pair 1 per pair 1 per pair 1 per pair 1 per pair	Launch Lab: Crystalline Columns, p. 47
Chapter 2, Section 2.1	polystyrene (StyrofoamTM) balls of different sizes, or gumdrops in two different sizes or colours toothpicks	44 × 10 = 440 large or 54 × 10 = 540 small for the class 1000	Investigation 2.A: Building Ionic Crystals, p. 51
Chapter 2, Section 2.1	100 mL beaker straw protractor hard, flat surface, at least 10 cm x 10 cm 15 cm ruler soap solution (mixture of 80 mL distilled water, 15 mL dish soap, and 5 mL glycerin) paper towels	1 per pair 1 per student 1 per pair 1 per pair 40 (2 per pair) 20 × 100 mL = 2000 mL 2 rolls	Investigation 2.B: Soap Bubble Molecules, pp. 56–57
Chapter 2, Section 2.2	two round balloons string marker	40 (2 per pair) 30 m 1 per pair	Investigation 2.C: Dipole Balloons, p. 64
Chapter 2, Section 2.2	500 mL beaker shallow dishes sewing needles small glasses 150 mL beaker acetate strip cotton cloth vinyl strip wool cloth vegetable oil water ethanol pepper liquid dish detergent	1 per pair 2 per pair 1 per pair 2 per pair 1 per pair 1 per pair 1 per pair 1 per pair 1 per pair 125 mL per pair 2L 2 L 5 shakers 5 small bottles	Investigation 2.D: Investigating the Properties of Water, pp. 66–67

Chapter, Section	**Item Description**	**Suggested Quantity** (assume 40 in class)	**Text Activity**
Chapter 2, Section 2.3	100 mL beaker and stirring rod	1 per group	Investigation 2.E: Properties of Substances, p. 78
	metal plate (iron or aluminium)	1 per group	
	conductivity tester	1 per group	
	candle	1 per group	
	Bunsen burner	1 per group	
	match, lighter, or striker	5 to be shared	
	ring stand with clamp	1 per group	
	timer	1 per group	
	tongs	1 per group	
	five unknown solids (total quantity depends on size of lab groups: these quantities assume partners)	(a) 20 solid chunks of each unknown solid for hardness test (b) 2 g in total of each solid for solubility test and melting point	
	distilled water	375 mL	

CHAPTER 1 CHEMICAL BONDING

Curriculum Correlation

General Outcome 1: Students will describe the role of modeling, evidence and theory used in explaining and understanding the structure, chemical bonding and properties of ionic compounds.

	Student Textbook	Assessment Options
Outcomes for Knowledge		
20–A1.1k recall principles for assigning names to ionic compounds	Unit Preparation, pp. 10-13 Launch Lab: Chemistry Recall, Chapter 1, p. 15 Binary Ionic Compounds, Unit 1 Preparation, p. 10	Launch Lab Analysis: 1, Chapter 1, p. 5 Section 1.1, Questions for Comprehension: 9, p. 34 Section 1.1 Review, 7–11, p. 35 Section 1.2 Review: 8, 10, p. 42 Chapter 1 Test BLM 1.0.7 Simple Binary Ionic Quiz BLM 1.0.8 Stock System Binary Ionic Compound Nomenclature and Formulas Quiz BLM 1.0.9 Mixed Ionic Compound Nomenclature and Formulas Quiz Unit 1 Review: 15, 16, pp. 86–89
20–A1.2k explain why formulas for ionic compounds refer to the simplest whole-number ratio of ions that result in a net charge of zero	Forming Ionic Bonds, Section 1.1, p. 20. 21 Using the Charge to Determine the Chemical Formula of an Ionic Compound, Section 1.2, p. 21	Section 1.1, Questions for Comprehension: 7, p. 21 Chapter 1 Review: 26, p. 44 Chapter 1 Test Unit 1 Review: 2–4, 13, 14, 40, 42, pp. 86–89
20–A1.3k define valence electron, electronegativity, ionic bond and intramolecular force	Unit 1 Preparation, pp. 8-9 Forming and Representing Compounds, Section 1.1, pp. 18-20, 23-29 Electronegativity, Section 1.2, p. 36 Size and Electronegativity, Section 1.2, p. 37 The Nature of Chemical Bonds, Section 1.2, pp. 39-42, 63	Section 1.1 Review: 2, 7, p. 35 Section 1.2 Review: 1, 3–5, 7, 8, 10, p. 42 Chapter 1 Review: 8, 19-21, p. 44-45 Chapter 1 Test Unit 1 Review: 1, 2, 4, 5, 7–9, 12, 17, 20, 26–28, 41, 45, 50, 52, 54, pp. 86–89
20–A1.4k use the periodic table and electron dot diagrams to support and explain ionic bonding theory	Forming Bonds, Unit 1 Preparation, p. 8–9 Forming and Representing Compounds, Section 1.1, pp. 19-20, 18-31 Electronegativity, Section 1.2, p. 36 Size and Electronegativity, Section 1.2, p. 37	Questions for Comprehension: 1–5, Section 1.1, p. 19 Section 1.1 Review: 1, 3–5, p. 35 Section 1.2 Review: 7, p. 42 Chapter 1 Review: 23, 26, 28, 34, p. 44-45 Unit 1 Review: 1, 2, 4, 7, 12–14, 26, 41, 51, 52, 57 , pp. 86–89

	Student Textbook	Assessment Options
Skill Outcomes (Focus on scientific inquiry)		
20–A1.5k explain how an ionic bond results from the simultaneous attraction of oppositely charged ions	Forming Bonds, Unit 1 Preparation, p. 9 Forming and Representing Compounds, Section 1.1, pp. 18-31	Section 1.1 Review: 3, 7, 9, p. 35 Section 1.2 Review: 2, 3, 6, 8, 10, p. 42 Chapter 1 Review: 9, 10, 14–17, p. 36 Chapter 1 Test Unit 1 Review: 4, 9, 21, 26, 27, 28, 32, 39, pp. 86–89
20–A1.6k explain that ionic compounds form lattices and that these structures relate to their properties, e.g., *melting point, dissolving, reactivity.*	Sample Problem: Drawing Lewis Structures, p. 15 Connections: Ionic Liquids, Section 1.1, p. 22 Three-Dimensional Structures, Section 2.1, pp. 48-50	Connections: Ionic Liquids, Section 1.1, p. 22 Chapter 1 Test Unit 1 Review: 3, 5, 6, 8, 10, 11, 30, 37–39, pp. 86–89
Outcomes for Science, Technology and Society (Emphasis on the nature of science)		
20–A1.1sts explain that the goal of science is knowledge about the natural world by ■ *identifying everyday processes and products in which ionic compounds are significant, such as in the composition of household products and foods and in life processes*	Clues in Naturally Occurring Substances, Section 1.1, p. 16 Connections: Ionic Liquids, Section 1.1, p. 22 Bonding in Biological Molecules, Section 2.2, pp. 69-70 Connections: Mane Products, Section 2.3, p. 79	Connections: Ionic Liquids: 1, 3, Section 1.1, p. 22 Unit 1 Review: 62–65, pp. 86–89
20–A1.2sts explain that scientific knowledge and theories develop through hypotheses, collection of evidence through experimentation and the ability to provide explanations by ■ *describing how an understanding of electronegativity contributes to the knowledge of relative bond strength, melting points and boiling points of ionic substances*	Clues in Naturally Occurring Substances, Section 1.1, p. 16 Connections: Ionic Liquids, Section 1.1, p. 22	Connections: Ionic Liquids: 1, 3, Section 1.1, p. 22 Chapter 1 Review: 33-35, p. 45 Unit 1 Review: 62–65, pp. 86–89
20–A1.3sts explain that scientific knowledge may lead to the development of new technologies and that new technologies may lead to scientific discovery by ■ *explaining how scientific research and technology interact in the production and distribution of beneficial materials, including semiconductors, ceramics and composite materials.*	Chapter 1 Opener, p. 14 Connections: Ionic Liquids, Section 1.1, p. 22 Career Focus: Ask a Nanotechnologist, p. 84-85	Career Focus: Ask a Nanotechnologist: 1-4, p. 84-85 Unit 1 Review: 62–65, pp. 86–89

	Student Textbook	Assessment Options
Skill Outcomes (Focus on scientific inquiry)		
Initiating and Planning		
20–A1.1s ask questions about observed relationships and plan investigations of questions, ideas, problems and issues by ■ designing an investigation to determine the properties of ionic compounds (solubility, conductivity, melting point) ■ describing procedures for safe handling, storage and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information ■ *researching the question "Should all scientific research have a practical application?"* ■ *designing an experiment to explore the formation of ionic compounds.*	Safety in Your Chemistry Laboratory and Classroom, p. xiv Connections: Ionic Liquids, Section 1.1, p. 22 Investigation 1.A: Modelling Molecules, Section 1.1, p. 30	Questions for Comprehension: 4, Unit 1 Preparation, p. 19 Connections: Ionic Liquids: 3, Section 1.1, p. 22 Investigation 1.A: Modelling Molecules: 3, Section 1.1, p. 30 Section 1.1 Review: 1, p. 35 Chapter 1 Review: 26, p. 44 Unit 1 Review: 45, 62–65, pp. 86–89
Performing and Recording		
20–A1.2s conduct investigations into relationships among observable variables and use a broad range of tools and techniques to gather and record data and information by ■ drawing electron dot diagrams and building models of ionic solids ■ *performing an investigation to illustrate properties of ionic compounds* ■ *using the periodic table to make predictions about bonding and nomenclature* ■ *using model building software to collect and integrate information on the structure of ionic crystals.*	Forming Ionic Bonds, Section 1.1, pp. 19-21 Electronegativity, Section 1.2, p. 36 Size and Electronegativity, Section 1.2, p. 37 Bond Type and Electronegativity, Section 1.2, p. 39 Investigation 2.A: Building Ionic Crystals, Section 2.1, p. 51	Section 1.2, Questions for Comprehension: 11, 12, p. 41 Investigation 2.A: Building Ionic Crystals, Analysis 1, Section 2.1, p. 51 Unit 1 Review: 45, pp. 86–89
Analyzing and Interpreting		
20–A1.3s analyze data and apply mathematical and conceptual models to develop and assess possible solutions by ■ *analyzing experimental data to determine the properties of ionic compounds* ■ *using data from various sources to predict the strength of bonds between ions.*	Connections: Ionic Liquids, Section 1.1, p. 22 The Nature of Chemical Bonds, Section 1.2, p. 36-42 Investigation 2.E: Properties of Substances, Section 2.3, p. 78	Connections: Ionic Liquids, Section 1.1, p. 22 Section 1.1 Review: 7-10, p. 35 Unit 1 Review: 8, 64, pp. 86–89
Communication and Teamwork		
20–A1.4s work as members of a team in addressing problems and apply the skills and conventions of science in communicating information and ideas and in assessing results by ■ critically analyzing models of ionic compounds built by others.	Investigation 2.A: Building Ionic Crystals, Section 2.1, p. 51	Investigation 2.A: Building Ionic Crystals, Analysis: 2, Section 2.1, p. 51

Chapter 1

Chemical Bonding

Student Textbook pages 14–45

Chapter Concepts

Section 1.1 Forming and Representing Compounds

- When atoms form bonds, they tend to gain, lose, or share electrons to achieve a filled outer energy level.
- Ionic bonds form when metal atoms lose electrons and non-metal atoms gain electrons. The resulting positive and negative ions attract each other by an electrostatic force.
- Covalent bonds form when non-metal atoms achieve a filled outer energy level by sharing electrons.
- Names of ionic compounds indicate which elements make up the compound.
- Names of some simple molecular compounds indicate the number of each type of atom in the molecule.
- Structural formulas show the bonds between atoms in a molecule.

Section 1.2 The Nature of Chemical Bonds

- Electronegativity is a measure of the relative ability of an element's atoms to attract electrons.
- Smaller atoms tend to have higher electronegativities.
- The difference in the electronegativities between two elements determines the type of bond that will form between atoms of those elements.

Common Misconceptions

- Students may think that gravitational forces are responsible for attractions between atoms and molecules. Point out that, due to the tiny mass of an atom, the effect of gravitational force is negligible. Electrical forces determine the physical and chemical properties of substances.
- Students may believe that atoms form bonds to satisfy the octet rule or that atoms borrow and lend electrons to satisfy the octet rule. However, the octet rule is descriptive of what happens when atoms combine; it is not a rule for atoms to follow. The reason the octet rule describes what is occurring has to do with energy considerations of filled and unfilled orbitals within electron shells.
- Students may not be aware that while a chemical bond can be identified as ionic, covalent, or metallic, all chemical bonds involve the electrostatic attraction between the positively charged nucleus of one atom and the valence electron(s) of the atom(s) to which it is bonded. This is the basis of all intramolecular bonding.
- Metallic bonding refers to the forces that exist between atoms in a pure metal or in a metal alloy. They do not involve the sharing/donation of electrons as in covalent/ionic bonding. Metallic bonding is used to explain the principal physical macroscopic properties of metals: malleability, ductility, excellent electrical conductivity, and heat conductivity.
- Most students understand the differences in the types of chemical bonds, but often miss the features that these bonds may have in common.
- Students may feel that the division between an ionic bond and a covalent bond is a sharp one. In fact, as ΔEN increases from 0 to 3.3, bonds become gradually more ionic and less covalent in character. So that there is a guide in place to decide whether a bond is ionic or polar covalent, chemists have decided to designate $\Delta EN = 1.7$ as the cutoff between ionic bonds and polar covalent bonds. It is important to realize that in practice there is no sharp dividing line between an ionic bond and a polar covalent bond at that point; the spectrum of ΔEN values is smooth. Have students examine Figure 1.22 on page 40 of the student textbook and ensure students realize that the cutoff at $\Delta EN = 1.7$ is there for classification purposes, and not because of a significant, experimentally observed change in properties of compounds at that point. The same arguments can be made for the designation of $\Delta EN = 0.5$ to distinguish between a polar covalent bond and slightly polar covalent bond.
- Students may think that electronegativity is a property of a single atom, since it has its own trend in the periodic table. Stress that, unlike atomic size and some other properties of isolated atoms, electronegativity is a property of an atom in a chemical bond. It is a measure of the tendency for an atom to attract electrons of another atom in a bonded situation, rather than an inherent property of the atom.

Helpful Resources

Books and Journal Articles

- Birk, James P.; Kurtz, Martha J., "Effect of Experience on Retention and Elimination of Misconceptions about Molecular Structure and Bonding," *Journal of Chemical Education,* 1999 (76) 124.
- Eberhart, Mark, *Why Things Break Apart: Understanding the World by the Way It Comes Apart,* Harmony Books, New York, 2003 ISBN: 1-4000-4768-9.
- Graham, Sarah, "Researchers Explain the Transformation of Water into Ice," *Scientific American,* April 2002.
- Lide, David R., CRC *Handbook of Chemistry and Physics,* 83rd Edition. 2004.
- Purser, Gordon H., "Lewis Structures Are Models for Predicting Molecular Structure, Not Electronic Structure," *Journal of Chemical Education,* 1999 (76) 1013.

Web Sites

Web links related to chemical bonding can be found at **www.albertachemistry.ca**. Go to the Online Learning Centre, and log on to the Instructor Edition. Choose Teacher Web Links.

List of BLMs

Blackline masters (BLMs) have been prepared to support the material in this chapter. The BLMs are either for assessment (AST); use as overheads (OH); use as handouts (HAND), in particular to support activities; or to supply answers (ANS) for assessment or handouts. The BLMs are in digital form, stored on the CD-ROM that accompanies this Teacher's Resource or on the web site at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs. The BLMs can be modified to suit the needs of your students.

Number (Type) Title

1.0.1 (HAND) Launch Lab: Chemistry Recall
1.0.1A (ANS) Launch Lab Answer Key
1.0.2 (OH) Electrons in Principle Energy Levels (first 20 elements)
1.0.3 (OH) Period 3 Electron Dot Diagrams
1.0.4 (OH) Electron Dot Diagrams: Formation of Ionic Compounds
1.0.5 (OH) Lewis Dot Diagrams for Hydrogen and Chlorine
1.0.6 (OH) Polyatomic Ion Table
1.0.7 (HAND/AST) Simple Binary Ionic Quiz
1.0.7A (ANS) Simple Binary Ionic Quiz Answer Key
1.0.8 (HAND/AST) Stock System Binary Ionic Compound Nomenclature and Formulas Quiz
1.0.8A (ANS) Stock System Binary Ionic Compound Nomenclature and Formulas Quiz Answer Key
1.0.9 (HAND/AST) Mixed Ionic Compound Nomenclature and Formulas Quiz
1.0.9A (ANS) Mixed Ionic Compound Nomenclature and Formulas Quiz Answer Key
1.0.10 (HAND/AST) Binary Molecular Compound Nomenclature and Formulas Quiz
1.0.10A (ANS) Binary Molecular Compound Nomenclature and Formulas Quiz Answer Key

1.1.1 (OH) Electron Dot Diagrams for Multi Ion Ionic Compounds
1.1.2 (OH) Lewis Structure of Methane and Water
1.1.3 (OH) Lewis Structures of Double and Triple Bonds
1.1.4 (OH) Rules for Drawing Lewis Structures
1.1.5 (HAND) Thought Lab 1.1: Lewis Structures
1.1.5A (ANS) Thought Lab 1.1: Lewis Structures Answer Key
1.1.6 (HAND) Investigation 1.A: Modelling Molecules
1.1.6A (ANS) Investigation 1.A: Modelling Molecules Answer Key
1.1.7 (OH) Metallic Bonding in Magnesium

1.2.1 (OH) Electronegativity of Elements
1.2.2 (OH) 3-D Periodic Table Trends in Electronegativities and Atomic Sizes
1.2.3 (OH) Ion Sizes and Attractive Forces
1.2.4 (OH) Bonding Spectrum: Covalent to Ionic
1.2.5 (OH) Bond Character and Electronegative Differences
1.3.1 (AST) Chapter 1 Test
1.3.1A (ANS) Chapter 1 Test Answer Key

Using the Chapter 1 Opener

Student Textbook pages 14–15

Teaching Strategies

- Have students review the periodic table. Have them relate properties of substances to the period number of each element, the number of electrons, and the energy levels the electrons occupy.
- Have students brainstorm as many molecular compounds (containing only non-metals) as possible. Have them write the formulas for these compounds, explain how these are named, and describe some of their properties.
- For fun, devise a game similar to the millionaire game. Write review questions, such as identifying symbols of elements, defining terms, or explaining processes on slips of paper. Form teams and draw the questions out of a box. Respondents may answer on their own, ask a friend, or ask the members of their teams. Keep track of team points.

Launch Lab Chemistry Recall

Student Textbook page 15

Purpose

The purpose of the activity is to help students recall concepts pertaining to elements, compounds, and bonding. Using symbols of chemicals written on small pieces of paper and drawn two at a time, students will attempt to form compounds. They will analyze whether the combination is possible, write the names and formulas of possible compounds, and identify them as ionic, covalent, or neither.

Outcomes

20–A1.1k

Advance Preparation

When to Begin	What to Do
1 week before	■ Ensure that there is one sheet of paper available for each partner pair.
1 day before	■ Decide how you will divide the class. ■ Photocopy **BLM 1.0.1 (HAND) Launch Lab: Chemistry Recall.**

Materials
■ scissors

Time Required

30 minutes

Helpful Tips

- Use **BLM 1.0.1 (HAND) Launch Lab: Chemistry Recall** to support this activity. Remove sections as appropriate to meet the needs of the students in your class. The answers to questions are on **BLM 1.0.1A (ANS) Launch Lab: Chemistry Recall Answer Key**.
- Provide the following bonus assignment: Ask students how many compounds they can make from their set of 25 symbols.

Answers to Analysis Questions

1. Students should recall the names of many of the symbols without looking at the periodic table. This question is for self-evaluation and to help students get into the mind-set of actively seeking to restore and enhance skills they started to acquire in *Science 10*.
2. Students should be able to identify whether some of the elements drawn will combine to form a compound, and whether the compound is ionic, covalent, or neither. Again, the purpose of this question is for self-evaluation. Reward answering the question about the review needed based on the results, instead of evaluating whether the answers were correct.
4. Reward creativity and extra effort as well as a workable game plan. Encourage students to have some fun with this one.

Assessment Options

- Assess students' participation and whether the partners discussed the results with each other. You may want to use parts of Assessment Checklist 2, 3, 4, and/or 5 found on the Teacher's Resource CD-ROM and in Appendix A.
- Have students hand in their answers to Analysis Questions 1 and 2.
- Assign Question 4 as a hand-in assignment/competition. Choose one or two and award prizes or bonus marks. Later, when there is time, play one of the games as an activity.

1.1 Forming and Representing Compounds

Student Textbook pages 16–35

Section Outcomes

Students will:

- explain how an ionic bond results from the simultaneous attraction of oppositely charged ions
- relate electron pairing to multiple and covalent bonds
- draw electron dot diagrams of atoms and molecules
- recall principles for naming ionic and molecular compounds

Key Terms

electron pairs
unpaired electrons
bonding pair
lone pairs
main group elements
molecular compounds
Lewis structures
single bond
double bond
triple bond
structural formula
metallic bonding
delocalized
nomenclature

Chemistry Background

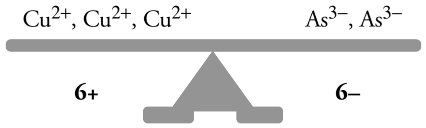
- Students are directed to focus on naturally occurring substances to find clues to the forces holding atoms to each other. Metallic ores and metals are examples. Metals are the most numerous of the elements. The periodic table, found in Appendix C of the student textbook, has the metals on the left-hand side of the dark staircase line starting under Group 13. Students should be familiar with some of the metals and with metallic properties from *Science 9* and *10*.
- Components of the atmosphere are examples of non-metallic elements. The noble gases, found in Group 18 at the extreme right-hand side of the periodic table, are distinct because they are never found in naturally occurring compounds. In 1962, Neil Bartlett of the University of British Columbia produced the first compound of inert gases. It was a compound of xenon, platinum, and fluorine. Bartlett was working alone on a Friday night at the time of the discovery. He was so excited that he rushed around to share this great news, but no one else was there. Bartlett was initially assisted by graduate student Derek H. Lohmann, who helped in the work that made Bartlett try this experiment.
- Students will have encountered the octet rule (sometimes referred to as the stable octet rule or rule of eight) in *Science 10*. The electrons in the outermost energy level of an atom are called valence electrons. When an atom has eight valence electrons, it usually will not bond with other atoms. Exceptions are hydrogen, helium, lithium, and beryllium, which are stable with only two electrons in their single energy level.
- The octet rule and the ionic versus covalent separation are constructs to help humans make sense of data. They are attempts to describe nature approximately through useful generalizations and categories, and are not rules nature follows explicitly. Exceptions and gradations of properties are part of the world we are trying to describe. Things are more complicated than presented here, but this is a strong framework on which to build further knowledge.
- All electrons are attracted by the nearest positive nucleus. When a metal atom gives up an electron to a non-metal

atom, it does not willingly "donate" the electron. The non-metal atom attracts the electron more strongly, as measured by the non-metal atom's electronegativity, and so pulls the electron away from the metal atom, producing the negative and positive ions.

- Some transition metals have a variable number of valence electrons because the energies of the electrons in a lower energy level are so close to that of the valence electrons that one or more electrons from the lower energy level can be "promoted" to the outermost energy level without significantly lowering the bond energy. The change in potential energy determines whether an electron will be "promoted" in this way.
- For ionic compounds, the formula represents the simplest ratio of ions, as if each single crystal element were a molecule. Since crystals differ in size, molecular formulas are meaningless. For even a very small crystal, indicating exact numbers of ions would give unmanageably large subscripts.
- The term "resonance structures" may seem to indicate that the electron structure resonates between forms, or switches from one form to the other. This is not true. The bond is somewhere in between a single bond and a double bond, with properties approximately equivalent to a bond and a half.
- The free electrons in metallic bonding are the valence electrons only.
- Compounds are named using IUPAC (International Union of Pure and Applied Chemistry) rules. The exact rules are also given in the *CRC Handbook of Chemistry and Physics.* Some molecular compounds have common names that do not follow IUPAC rules. These names were in use before the IUPAC system was established and are commonly accepted (for example, sucrose, benzene).

Teaching Strategies

- Build on the self-discovery from the Launch Lab. Emphasize students' strengths and encourage their discoveries so they can refine the knowledge they acquired previously.
- Ensure students understand that the models represent certain characteristics of the real substances, but do not illustrate the actual appearance of the structures. (Carbon atoms are not really black solid spheres and bonds do not look like sticks or springs, etc.)
- To help students review and learn names of polyatomic ions, suggest that they first learn the name of a common polyatomic ion ending in -ate, ClO_3^-. Point out that one less oxygen atom results in a name ending in -ite; two less oxygen atoms result in a name beginning with hypo- and ending in -ite.
- For students having difficulty balancing charges for multiple ions, set up a balance apparatus to give them another way of looking at why the subscripts are chosen. For example, for Cu_3As_2, the positive charges $(3 \times 2^+)$ balance the negative charges $(2 \times 3^-)$.



- When you use Figure 1.9 (student textbook page 21), be sure to emphasize that if the formula is not the simplest whole number ratio, it must be reduced. Many students simply transpose superscripts to subscripts, without understanding the process of balancing the charges.
- Keep in mind that most of this section is a review of skills studied in *Science 10*. For many students, the treatment there will have been too fast for them to have developed the ability to be consistently correct. This is the time to make sure they can. Quickie-quizzes of five to 10 questions of slightly increasing difficulty, at the beginning of class several days in a row, help many students focus on perfecting their skills at naming and writing formulas, and build confidence. Quizzes also highlight deficiencies that need to be addressed. Students need to be made aware that there will be a series of quizzes and told what skills will be tested on the next quiz.
- For students to become competent in naming compounds, it is important that they practise. To make the amount of practice that is required more interesting, one or more of the following strategies could be used:
  - Divide the class into pairs for written exercises. Try to pair the more competent students with those having difficulty in order to encourage peer coaching.
  - Divide the class into pairs for practice. Send one of each pair to the chalkboard and have the other work at their desk. Give them chemical formulas of ionic and molecular substances and ask them to name the compounds and draw electron dot diagrams of each. Part-way through the exercise switch those at the chalkboard with those in their seats. Rotate and switch partners for the next set.
  - Make a game of learning compound names. For example, devise a game similar to bingo, but print the cards with chemical formulas instead of numbers. Read out names of compounds while students mark off the corresponding chemical formulas on their cards.
- A number of overhead masters have been prepared for this section. You will find them with the Chapter 1 BLMs on the CD-ROM that accompanies this Teacher's Resource or at **www.albertachemistry.ca**, Online Learing Centre, Instructor Edition, BLMs.

  Number (Type) Title
  1.1.1 (OH) Electron Dot Diagrams for Multi Ion Ionic Compounds
  1.1.2 (OH) Lewis Structure of Methane and Water
  1.1.3 (OH) Lewis Structures of Double and Triple Bonds
  1.1.4 (OH) Rules for Drawing Lewis Structures
  1.1.7 (OH) Metallic Bonding in Magnesium
- Use the Connections (Nature of Science) feature about ionic liquids on page 22 of the student textbook as

extension material for students who have mastered the concept of ionic compounds.

Chemistry File: Web Link

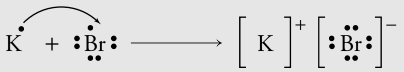
Student Textbook page 17

Helium has a low solubility in blood, which is made of mostly water. Helium is used in deep-sea diving to decrease the risk of decompression sickness that can result from high-pressure oxygen and nitrogen. Argon is used as a shield of inert gas to prevent welded materials from oxidizing.

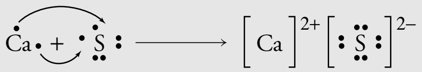
Answers to Questions for Comprehension

Student Textbook page 19

Q1. (a)



(b)

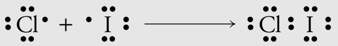


(c) A compound is not possible because both atoms need to lose electrons.

(d) Only a covalent (molecular) compound is possible because both atoms need to gain (or share) electrons.

Q2. (a) A molecular compound will not form because neon has a full octet.

(b)



(c) A molecular compound will not form because lithium's electronegativity is too low.

(d)

Q3. The octet rule states that no more than eight electrons can exist in the outer energy level of any atom, and that when bonds form, atoms gain, lose, or share valence electrons in such a way that they create a filled outer energy level of electrons. It appears that this filled outermost energy level is especially stable.

Q4. 1. Draw the symbol for the nitrogen atom.

2. Determine the number of valence electrons in a nitrogen atom. (5)

3. Begin on any of the four sides around the symbol, and place one dot on each side until all four sides have one dot.

4. Pair the last electron with an existing single.

Q5. A valence electron is an electron in the outermost energy level or shell of an atom.

Chemistry File: Try This

Student Textbook page 20

To calculate the total number of electrons in a given energy level, use the formula $2n^2$. For example, the third energy level (n=3) can hold $2(3)^2 = 18$ electrons. This works because the third energy level has three sublevels. The s sublevel holds 2 electrons, the p sublevel holds 6 electrons, and the d sublevel holds 10 electrons, so $2 + 6 + 10 = 18$.

Chemistry File: Try This

Student Textbook page 21

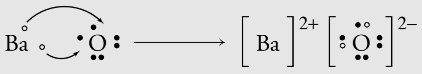
In general, the greater the horizontal distance between two elements on the periodic table, the more likely they are to be involved in ionic bonding. Metals, which are on the left side of the table, usually have three or fewer electrons in their outer energy levels and can achieve a stable octet by losing these electrons, and nonmetals, on the right side of the table, have four or more electrons in their outer energy level and can achieve a stable octet by gaining four or more electrons.

Answers to Questions for Comprehension

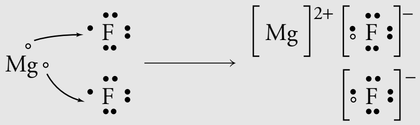
Student Textbook page 21

Q6. (a) KCl

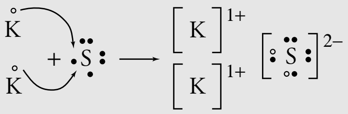
(b) BaO



(c) MgF_2



(d) K_2S



Q7. (a) ZnO

(b) FeS

(c) MgI_2

(d) $CoCl_3$

Connections (Nature of Science): Ionic Liquids

Student Textbook page 22

Answers to Questions

1. Ionic liquids conduct electricity because the cations and anions can migrate in opposite directions within the liquid. This results in conductivity because charge is transferred. Possible applications include advanced batteries and super capacitors, and the refining of sodium

metal by electricity. The following web site has a concept map showing applications of ionic liquids: **http://quill.qub.ac.uk/map/**.

2. Green chemistry is a movement to create and use environmentally friendly cleaning products, paint, gardening products, and other chemical processes, and to prevent pollution on a molecular level. In taking their survey, students should first ask what their friends and family know about green chemistry. Before asking if they feel green chemistry is important, students may have to clarify with their subjects the meaning of green chemistry.
3. Because ionic liquids dissolve petroleum-based liquids, ionic liquids can replace industrial solvents, which are a major source of air pollution in manufacturing everything from plastics to drugs. It would be important to establish that the ionic liquid chosen would not be toxic if it migrated to water supplies.

Answers to Practice Problems 1–6

Student Textbook page 25

For full solutions, visit **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, Full Solutions.

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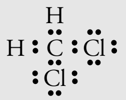

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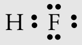

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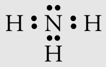

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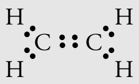

Answers to Practice Problems 7–17

Student Textbook page 27

For full solutions, visit **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, Full Solutions.

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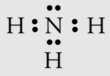

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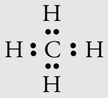

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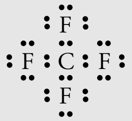

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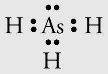

11.


12.


13.


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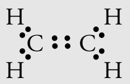

16.


17.


Thought Lab 1.1: Lewis Structures

Student Textbook page 27

Purpose

Students attempt to draw Lewis structures of a variety of larger molecules that do not have a central atom. Working with a partner, they then objectively evaluate their models and try to come to a consensus on a correct model.

Outcomes

20–A2.4k

Advance Preparation

None

Time Required

20–30 minutes

Helpful Tips

- Use **BLM 1.1.5 (HAND) Thought Lab 1.1: Lewis Structures** to support this activity. Remove sections as appropriate to meet the needs of the students in your class.

Answers to Questions

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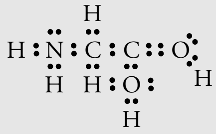

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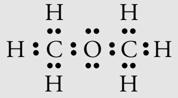

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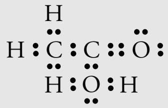

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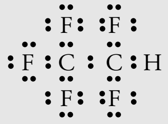

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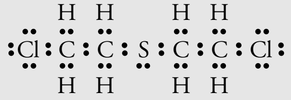

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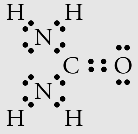

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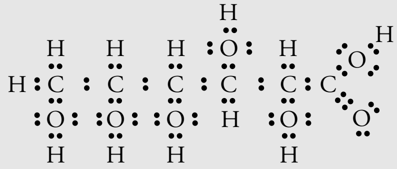


Assessment Options

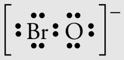
- Collect and assess answers to questions.

Answers to Practice Problems 18–24

Student Textbook page 29

For full solutions, visit **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, Full Solutions.

18.



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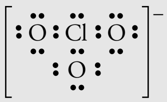

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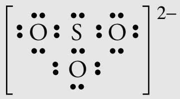

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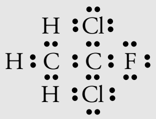
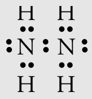


23. (a)



23. (b)

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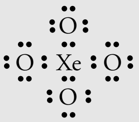


Investigation 1.A: Modelling Molecules

Student Textbook page 30

Purpose

Students build models of molecules to match Lewis structures and to fit some of the properties that have been determined by experimentation. Students will name and draw Lewis structures for 10 molecular compounds to illustrate the shared electrons of the atoms making up the substances. They will then build models of these structures using a prepared molecular model kit. The purpose is not just to practise these skills, but also to acquire a better sense of the three-dimensional nature of the bonds and structures.

Outcomes

- 20–A1.6k
- 20–A1.2s
- 20–A1.4s
- 20–A2.4k

Advance Preparation

When to Begin	What to Do
2–3 weeks before	■ Ensure that molecular modelling kits are available.
1 day before	■ Collect the molecular modelling kits to distribute to the students.

Materials

- a molecular modelling kit for each student or group of students
- (optional) photocopy **BLM 1.1.6 (HAND) Investigation 1.A: Modelling Molecules** for each student

Time Required

45 minutes

Helpful Tips

- Use **BLM 1.1.6 (HAND) Investigation 1.A: Modelling Molecules** to support this activity. Remove sections as appropriate to meet the needs of the students in your class. The answers to questions are on **BLM 1.1.6A (ANS) Investigation 1.A: Modelling Molecules Answer Key.**
- Review drawing Lewis structures, emphasizing that all atoms in this investigation except hydrogen will need an octet of electrons available, and that hydrogen will need two electrons.
- Students tend to believe that the models are accurate representations, and you will need to be very clear that the models show relative positions of atoms, but that atoms are not solid, nor are they coloured, and bonds are not visible entities. Any colour coding is only to help identify which elements are where. Review the Common Misconceptions for Chapter 1 on page 10 of this Teacher's Resource.
- You may wish to assign students to their groups rather than letting them choose. Assign (or let students assign) roles to each member of the group, such as recorder, task manager, sketcher, constructor(s), namer, Lewis structure checker, selector of other molecules, and so on.
- Having different types of molecular model kits can bring on a discussion of relative merits of different styles, for example ball and spring versus ball and stick, and so on.
- If not enough kits are available, marshmallow and toothpick combinations work. Purchase coloured marshmallows in both large and small sizes. Students will have to wash up after using them. Tell them that they can eat some fresh ones when the work is done.

Answers to Analysis Questions

1. Ensure that students distinguish between differences caused by errors and differences in the view taken of the structure. (One being the upside-down version of the other, etc.)
2. The three-dimensional nature of the structures and the relative lengths and angles of bonds may be better represented by models than by Lewis structures.

Answer to Conclusion Question

3.

Strengths of using molecular kits	Limitations of using molecular kits	What can be deduced from the models	Features not deduced from the models
■ model is three-dimensional	■ model is not an accurate representation of molecular structure ■ students may develop a misconception about what atoms and bonds look like	■ general shape of the molecule ■ number and ratio of atoms in the molecule ■ direction and number of the bonds	■ actual appearance of molecules ■ position of non-valence electrons ■ strength of the bonds ■ actual size of the atoms

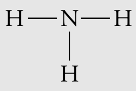
Assessment Options

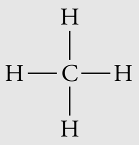
- Complete the relevant parts of Assessment Checklist 5: Learning Skills for each student during the investigation. The checklist is found on the Chemistry Teacher's Resource CD-ROM (see Appendix A).
- Have students complete Checklist 3: Performance Task Self-Assessment. Since some questions may not apply to this activity, specify to students which questions they are to answer.
- Collect and assess student responses to Analysis Questions 1 and 2 and Conclusion Question 3. Assess for thoughtfulness and completeness as well as accuracy.

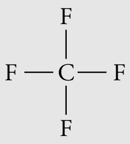
Answers to Questions for Comprehension

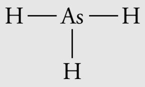
Student Textbook page 31

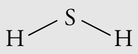
Q8.

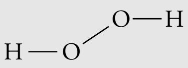
a)

$H-\underset{\displaystyle H}{N}-H$

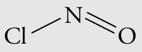
b)

$H-\underset{\displaystyle H}{\overset{\displaystyle H}{C}}-H$

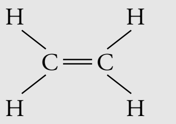
c)

$F-\underset{\displaystyle F}{\overset{\displaystyle F}{C}}-F$

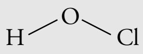
d)

$H-\underset{\displaystyle H}{As}-H$

e)


f)


g)


h)


i)


j) $H-C\equiv C-H$

Chemistry File: Web Link

Student Textbook page 32

An alloy is a combination of a metal and one or more other elements. Alloys are held together by metallic bonds.

Answers to Questions for Comprehension

Student Textbook page 34

Q9. (a) silicon dioxide
(b) cesium iodide
(c) sodium hydrogen phosphate
(d) dinitrogen pentoxide
(e) barium hydroxide
(f) nitrogen tribromide
(g) ammonium hydrogen phosphate
(h) hydrogen chlorate
(i) aluminium oxide
(j) magnesium nitrate
(k) potassium hypochlorite
(l) potassium permanganate

Section 1.1 Review Answers

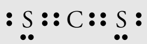
Student Textbook page 35

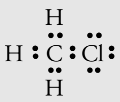
1. The noble gases are especially stable, and, with the exception of helium, each have eight valence electrons. This leads to the idea that when atoms have the same electron configuration as one of the noble gases, a stable octet, they will no longer react.

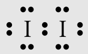
2. (a) Valence electrons are the electrons in the outermost energy level or shell of an atom.
(b) A lone pair is a pair of unbonded valence electrons in an atom.
(c) An unpaired electron is a valence electron that has no partner to form a pair. It is available to participate in bonding.
(d) An ionic bond is an electrostatic attraction between oppositely charged ions.

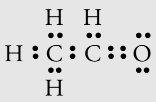
3. (a) 1 : 2 **(b)** 1 : 1 **(c)** 1 : 1 **(d)** 1 : 1

4. No dots signify that the atom has lost all of its valence electrons, and has become a positive ion. (Except in the case of hydrogen, the ion would still have electrons in filled lower energy levels.)

5. (a)


(b)


(c)


(d)


6 (a) The atoms contribute only 17 valence electrons, but the structure has 18.

or

The second carbon atom has 10 electrons surrounding it instead of eight.

(b) The hydrogen atom has eight electrons surrounding it instead of two.

(c) Both the oxygen atom and the carbon atom have 10 electrons surrounding them instead of eight.

or

The atoms only contribute 13 valence electrons, but the structure has 16.

7. (a) magnesium chloride **(b)** sodium oxide
(c) iron(III) chloride **(d)** copper(II) oxide
(e) zinc sulfide **(f)** aluminium bromide

8. (a) $FeCl_3$ **(b)** MgO **(c)** KBr
(d) Ca_3P_2 **(e)** MnS **(f)** BaI_2

9. (a) potassium chromate **(b)** ammonium nitrate
(c) sodium sulphate **(d)** strontium phosphate
(e) potassium nitrite **(f)** barium hypochlorite

10. (a) $NaHCO_3$ **(b)** $Na_2S_2O_3$
(c) $NaClO$ **(d)** $LiNO_2$
(e) $KMnO_4$ **(f)** NH_4Cl

11. (a) sulfur dioxide **(b)** carbon monoxide
(c) dinitrogen pentoxide **(d)** dichlorine monoxide

12. (a) H_2O **(b)** SO_3 **(c)** $SiCl_4$

13. (a) Atoms are represented by A, F, and G, since the number of electrons and protons are equal making a net charge of zero. Ions are represented by B, C, D, and E, since the number of electrons and protons are not equal.

(b) B^{3-}, C^{+}, D^{2+}, E^{-}

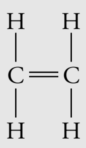
(c) A is B (boron), B is N^{3-} (nitride ion), C is K^{+} (potassium ion), D is Zn^{2+} (zinc ion), E is Br^{-} (bromide ion), F is B (boron, different isotope from A), G is F (fluorine).

14. (a) Potassium bromide is not a molecule, but an ionic compound. It would be appropriate to say "four formula units of potassium bromide."

(b) Since the ion in this compound contains hydrogen, the name of the compound is sodium hydrogen sulfate.

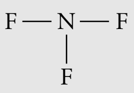
(c) Potassium nitrate is KNO_3. This compound, KNO_2, is potassium nitrite.

15. (a)



$$\begin{array}{ccc} H & & H \\ | & & | \\ C & = & C \\ | & & | \\ H & & H \end{array}$$

(b)



$$\begin{array}{ccccc} F & - & N & - & F \\ & & | & & \\ & & F & & \end{array}$$

(c) $H—C \equiv N$

1.2 The Nature of Chemical Bonds

Student Textbook pages 36–42

Section Outcomes

In this section, students will:

- define electronegativity
- describe bonding as a continuum

Key Terms

electronegativity
polar covalent bonds
bond dipoles
electronegativity difference
non-polar covalent bonds

Chemistry Background

- Electronegativity, the attraction of an atom for shared electrons, is a value determined experimentally or computed from theory. The scale is for comparison only, and relative values range from 0.7 to 4.0. There are no units for electronegativity because it is a relative measure and therefore dimensionless.
- In *Science 9* and *10*, the terms **ionic** and **covalent** were introduced to students as if they were mutually exclusive categories. This is oversimplification. Ionic and covalent are separate categories for naming purposes, but properties of bonds lie in a continuum between totally ionic and totally covalent.

Teaching Strategies

- The caption for Figure 1.23 on page 41 of the student textbook asks how dissolving hydrogen and chlorine in water affects the bonding. This could be used as an introduction to the concept that the bond in a substance like HCl (hydrochloric acid) will be altered by the polarity of the ions in water. Polarity affects the properties of the substance: hydrochloric acid in water conducts electricity and is acidic.
- Electronegativity is a very abstract concept and the nature of the word might suggest that as a characteristic it might have "negative" implications. Electronegativity is neither positive nor negative in terms of desirability. An analogy for electronegativity might be the ability of a pop star to attract fans. Emphasize that in this analogy, the ability to attract can be seen as "positive" from the pop star's point of view, perhaps "negative" from a parent's point of view, but the electronegativity of an element is neither a "good" nor a "bad" quality, it simply is.
- Ensure that students distinguish between the characteristics of the bond and whether the substance is a molecule. For example, in the case of NaCl, which is an ionic compound, you would not say "the partial positive charge is on the sodium end of the molecule." You *would* say "…the sodium end of the bond."

- A number of overhead masters have been prepared for this section. You will find them with the Chapter 1 BLMs on the CD-ROM that accompanies this Teacher's Resource or at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs.
  Number (Type) Title
  1.2.1 (OH) Electronegativity of Elements
  1.2.2 (OH) 3-D Periodic Table Trends in Electronegativities and Atomic Sizes
  1.2.3 (OH) Ion Sizes and Attractive Forces
  1.2.4 (OH) Bonding Spectrum: Covalent to Ionic
  1.2.5 (OH) Bond Character and Electronegative Differences

Chemistry File: Web Link

Student Textbook page 37

The first periodic table was created by Dmitri Mendeleev in the late 1860s. Mendeleev arranged the elements in order of increasing atomic mass, but he also paired the increasing atomic mass with recurring properties, such as atomic radius, chemical properties, electron configuration, electronegativity, and ionization energy. His periodic table actually left gaps for missing elements that were filled in later, when the elements were discovered.

Chemistry File: Web Link

Student Textbook page 38

Coulomb used a device called a torsion balance to measure the attraction between balls with varying charges at varying distances from each other.

Answers to Questions for Comprehension

Student Textbook page 41

Q10. **(a)** $\Delta EN = 1.43$, therefore the bond is polar covalent
(b) $\Delta EN = 0.40$, so the bond is slightly polar covalent
(c) $\Delta EN = 0$, so the bond is non-polar covalent
(d) $\Delta EN = 1.54$, so the bond is polar covalent
(e) $\Delta EN = 0.30$, so the bond is slightly polar covalent
(f) $\Delta EN = 3.05$, so the bond is mostly ionic
(g) $\Delta EN = 1.61$, so the bond is polar covalent
(h) $\Delta EN = 1.89$, so the bond is mostly ionic

Q11. **(a)** partial positive: carbon end; partial negative: fluorine end
(b) partial positive: nitrogen end; partial negative: oxygen end
(c) no partial positive or partial negative
(d) partial positive: copper end; partial negative: oxygen end
(e) partial positive: silicon end; partial negative: oxygen end
(f) partial positive: sodium end; partial negative: fluorine end
(g) partial positive: iron end; partial negative: oxygen end
(h) partial positive: manganese end; partial negative: oxygen end

Q12. **(a)** carbon bonded to sulfur ($\Delta EN = 0.03$), oxygen bonded to nitrogen ($\Delta EN = 0.40$), hydrogen bonded to chlorine ($\Delta EN = 0.96$), sodium bonded to chlorine ($\Delta EN = 2.23$).
(b) nitrogen bonded to nitrogen ($\Delta EN = 0$), carbon bonded to chlorine ($\Delta EN = 0.61$), phosphorus bonded to oxygen ($\Delta EN = 1.25$), magnesium bonded to chlorine ($\Delta EN = 1.85$).

Section 1.2 Review Answers

Student Textbook page 42

1. Electronegativity is the relative measure of an atom's ability to attract the shared electrons in a chemical bond.

2. The electrostatic force is the mutual attraction between negative and positive charges. In ionic bonds, there is an attraction between positive and negative ions; in covalent bonds, there is an attractive force between the shared electrons and the positive nuclei of the atoms.

3. As you look across any period on the periodic table from left to right, electronegativity increases as the atomic number increases until you reach the noble gases.

4. As you look down any group (except the noble gases) on the periodic table, electronegativity decreases.

5. There is an inverse relationship between the size of an atom and its electronegativity. As the number of protons in the nucleus increases, the attractive force on the electrons increases, pulling them closer to the nucleus. The size of an atom determines how close bonding electrons can get to the nucleus. Therefore, size affects how strongly the positively charged nucleus can attract those electrons in the bond. As size decreases across the period, electronegativity increases. As size increases down a group, electronegativity decreases.

6. There is no clear distinction between bond types. Bond types are viewed as parts of a "bonding continuum" and cannot be rigidly categorized.

7. In order of increasing attraction:
(a) N and O, $\Delta EN = 0.40$
(b) H and Cl, $\Delta EN = 0.96$
(c) Mn and O, $\Delta EN = 1.89$
(d) Ca and Cl, $\Delta EN = 2.16$

8. **(a)** Mostly ionic bonds are formed between atoms whose ΔEN is between 1.7 and 3.3.
(b) Polar covalent bonds are formed between atoms whose ΔEN lies between 0.5 and 1.7.

(c) Non-polar covalent bonds are formed between atoms whose ΔEN is between 0 and 0.5.

9. Potassium and fluorine would form a non-polar bond.

10. **(a)** mostly ionic
(b) non-polar covalent
(c) non-polar covalent
(d) mostly ionic

Chapter 1 Review Answers

Student Textbook pages 44–45

1. Noble gases are never found bonded in nature, so they are very stable. Atoms of noble gases also have their outermost energy level filled, with eight electrons for all but helium. This leads to the conclusion that in bonding, atoms can become more stable by gaining, sharing, or losing electrons to achieve an outermost energy level of eight electrons like a noble gas. This is the octet rule.

2. Hydrogen does not conform to the octet rule because its outermost energy level becomes filled with only two electrons. In order to achieve a filled outermost energy level, it needs to gain only one electron or share two. Helium also does not conform; it has two electrons in its outer shell.

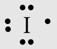
3. An electron dot diagram is a model that shows the valence electrons of the atoms surrounding the chemical symbol of the atom. The electrons are often arranged in pairs, shared electrons are shown between atoms, and the diagram also demonstrates how electrons form octets in compounds.

4. (a)

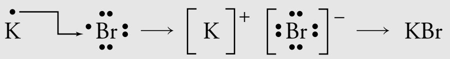
(b)

(c)

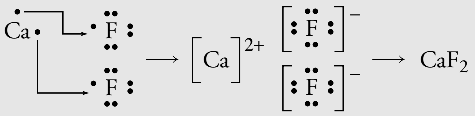
(d)



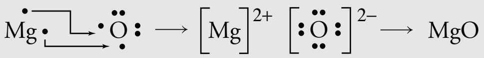
5. (a)



(b)



(c)



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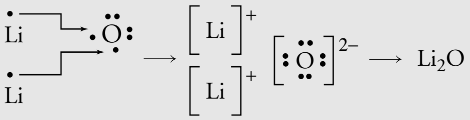


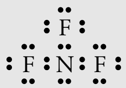
6. The number of electrons in non-metals that are involved in bonding corresponds to the number of single electrons in the electron dot diagram. Lone pairs are not involved in bonding.

7. (a)

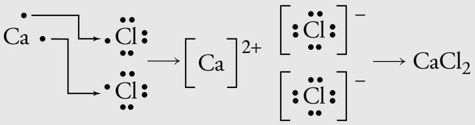
(b)

(c)

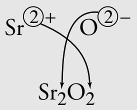
(d)



8.



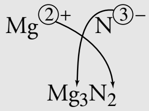
9. Strontium oxide



The lowest whole number ratio is 1:1 so the formula is:

SrO

Magnesium oxide



The lowest whole number ratio is 3:2 so the formula is:

Mg_3N_2

10. A lone pair is a pair of electrons in an atom's shell that are not involved in bonding. A bonding pair is a pair of electrons that are shared between two atoms to form a bond between the atoms.

11. Each oxygen atom contributes two electrons to form a double bond. In effect, the two oxygen atoms share four electrons, making a stable octet for each atom.

12. In a structural formula, one bonding line replaces the two dots of a bond in the Lewis structure. Two bonding lines replace the four dots of a double bond in a Lewis structure. Non-bonding (lone) electron pairs are not shown. Otherwise, the structure shown for the molecule remains the same.

13. Metallic bonds cannot be ionic since metals tend only to lose electrons, forming positive ions, so there are no negative ions with which to bond. If two metals shared all their valence electrons, there would still be less than eight in the outermost energy level, so the bonds cannot be covalent.

14. Valence electrons in metals are held loosely and are free to move from atom to atom. The electrostatic force between the positively charged metal ions and the cloud of negative electrons surrounding them create the metallic bonding. This is the free-electron model of metallic bonding.

15. **(a)** SnF_2 **(b)** $BaSO_4$ **(c)** HCN
(d) CsBr **(e)** $(NH_4)_2HPO_4$ **(f)** $NaIO_4$
(g) $KBrO_3$ **(i)** $Ca(ClO)_2$ **(j)** $Pb_3(PO_4)_2$
(h) NaCN

16. **(a)** hydrogen iodite **(b)** cesium fluoride
(c) sodium hydrogen sulfate **(d)** potassium dichromate
(e) potassium perchlorate **(f)** dinitrogen dichloride
(g) aluminium sulfate **(h)** iron(III) periodate

17. **(a)** ClF **(b)** P_2O_5
(c) SiS_2 **(d)** P_2Se_3

18. **(a)** carbon disulfide **(b)** iodine monochloride
(c) sulfur dioxide **(d)** dinitrogen trioxide
(e) phosphorus trichloride **(f)** selenium dibromide

19. Electronegativity applies to bonding or valence electrons.

20. Within a period, excluding noble gases, electronegativity increases from left to right. Down a group, electronegativity decreases from top to bottom.

21. **(a)** Li, La, Zn, Si, Br
(b) Cs, Y, Ga, P, Cl

22. In a non-polar bond, the sharing of electrons is equal between the atoms; in a polar bond, the sharing of electrons is unequal, with one atom attracting the electrons more strongly than the other.

23. **(a)** $\Delta EN = 1.79$, bond is ionic
(b) $\Delta EN = 1.35$, polar covalent
(c) $\Delta EN = 1.28$, polar covalent
(d) $\Delta EN = 0.40$, slightly polar covalent

24. **(a)** Mn and F, Mn and O, Mn and N
(b) Be and F, Be and Cl, Be and Br
(c) Ti and Cl; Fe and Cl; tie between Cu and Cl, and Hg and Cl; Ag and Cl

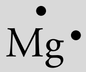
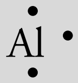
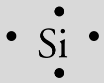
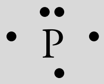
25. Since a polar covalent bond has some similarities with an ionic bond, and these similarities vary with the degree of polarity as measured by ΔEN, bonds form a continuum from non-polar covalent through polar covalent to ionic, with corresponding increases in percent ionic character and decreases in percent covalent character. Most bonds have some of each character.

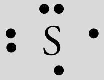
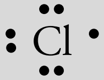
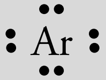
Answers to Applying Concepts Questions

26. Ensure that models demonstrate that in ionic compounds electrons are transferred and in molecular compounds electrons are shared. The idea that molecular compounds exist as discrete molecules but ionic compounds consist of innumerable formula units in a lattice would be a bonus. Assess and reward creativity, effectiveness, and artistic merit.

27. Since the elements Li, Be, B, C, N, O, F, and Ne are in order, the number of valence electrons corresponds to their position. Knowing that lithium is in the first place, it follows that lithium has one valence electron, beryllium has two, and the rest in order have three, four, five, six, seven, and eight.

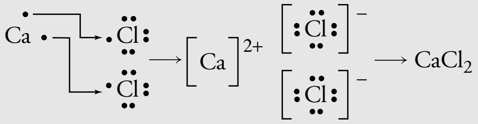
28. **(a)**

Group #	Example	Electron dot diagram	# valence electrons	# lone pairs	# unpaired electrons (number of bonds normally formed)	Nature of bonds formed (ionic, metallic, covalent)
1	Na, sodium		1	0	1	Ionic or metallic
2	Mg, magnesium		2	0	2	Ionic or metallic
13	Al, aluminium		3	0	3	Ionic or metallic
14	Si, silicon		4	0	4	Covalent
15	P, phosphorus		5	1	3	Covalent (nitrogen may form ionic bonds)

Group #	Example	Electron dot diagram	# valence electrons	# lone pairs	# unpaired electrons (number of bonds normally formed)	Nature of bonds formed (ionic, metallic, covalent)
16	S, sulfur		6	2	2	Ionic or covalent
17	Cl, chlorine		7	3	1	Ionic or covalent
18	Ar, argon		8	4	0	Usually does not form bonds

(b)
- The number of valence electrons increases from one to eight as you scan from Group 1 to Group 18.
- The number of electron pairs is zero for the first four groups and then increases from 1 to 4 for the last four groups.
- The number of unpaired electrons increases from 1 to 4 for the first four groups and then decreases from 3 to 0 for the last four groups.

29. An example might be as follows:



30. **(a)** C, carbon

(b) O, oxygen

(c) C, carbon

(d) N, nitrogen

(e) C, carbon

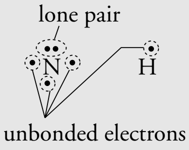
31. **(a)** Metallic bonds exist between atoms of iron in the sparkler.

(b) Covalent bonds exist between atoms of oxygen in air.

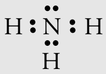
(c) Ionic bonds are formed between iron ions and oxide ions when the sparkler is burned.

32.

(a) & **(b)**



(c)



(d) The bonds formed between the hydrogen and nitrogen atoms are polar covalent bonds.

Answers to Making Connections Questions

ICT **33.** The multimedia presentation could be done using a video with sound bites and/or music, in a computer slide presentation, with charts and a tape recorder, as a rap music video with actions and demonstrations, and so on. Reward visual/auditory/emotional impact, as well as creativity and content.

34. Borax contains sodium tetraborate decahydrate, which is ionic. Dishwasher detergent contains calcium carbonate and sodium silicate, which are both ionic. Envirodesign Kitchen Cleaner^{TM} contains hydrogen anionic surfactants, so you could predict it will be molecular (anionic means "no ions"). Vinegar contains acetic acid, which is molecular. Most cleaners do not give the chemical contents, so students may have to predict from the cautions given or look at hair cleaning or personal care products, which often list ingredients. Epsom salt, $MgSO_4(s)$, is an ionic compound found in hard water. It is a solid crystal at room temperature and dissolves readily in water. It is made up of metallic magnesium ions joined to polyatomic ions composed of sulfur and oxygen. The bond must be strong because Epsom salt has a relatively high melting and boiling point and does not break down when stored for a long time. Coconut oil is a naturally occurring molecular compound made of carbon, hydrogen, and oxygen. It is a liquid at room temperature with a low melting temperature. The oil does break down over time, so the bonds are not as strong as those of Epsom salt.

35. A prediction could be "The greater the polarity of the bonds in a compound, the greater the solubility of the compound in water."

CHAPTER 2 DIVERSITY OF MATTER

Curriculum Correlation

General Outcome 2: Students will describe the role of modeling, evidence and theory used in explaining and understanding the structure, chemical bonding and properties of molecular substances.

	Student Textbook	Assessment Options
Outcomes for Knowledge		
20–A2.1k recall principles for assigning names to molecular substances	Covalently Bonded Compounds, Unit 1 Preparation, p. 10-11	Questions for Comprehension: 5, 6, p. 13 Section 1.1 Review: 10-12, p. 35 Chapter 1 Review: 16, 18, p. 44-45 Practice Problems: 1-10, Section 2.1, p. 56 Chapter 2 Test Unit 1 Review: 3, 13, 15, 16, pp. 86–89
20–A2.2k explain why formulas for molecular substances refer to the number of atoms of each constituent element	Covalently Bonded Compounds, Unit 1 Preparation, p. 10-11	Questions for Comprehension: 5, 6, p. 13 Chapter 2 Test Unit 1 Review: 4, 9, 11, 13, 31–34, 40, 45, 46, 50, 58, pp. 86–89
20–A2.3k relate electron pairing to multiple and covalent bonds	Chemical Bonds, Section 1.1, p. 18 Forming Covalent Bonds, Section 1.1, p. 23 Coordinate Covalent Bonds, Section 1.1, p. 28 VSEPR and the Structures of Molecular Compounds, Section 2.1, p. 52	Questions for Comprehension: 2, p. 19 Section 1.1 Review: 2, p. 35 Section 1.2 Review: 2, 7, p. 42 Chapter 2 Test Unit 1 Review: 1, 4, 9, 12, 13, 19, 20, 21, 23, 34, 49, 50, pp. 86–89
20–A2.4k draw electron dot diagrams of atoms and molecules, writing structural formulas for molecular substances and using Lewis structures to predict bonding in simple molecules	Electron Dot Diagrams, Unit 1 Preparation, p. 8 Clues in Naturally Occurring Substances, Section 1.1, p. 16-18 Forming Covalent Bonds, Section 1.1, p. 23 Sample Problem: Drawing Lewis Structures, Section 1.1, p. 24 Drawing Lewis Structures for Simple Molecules and Polyatomic Ions, Section 1.1, p. 25 Sample Problem: Drawing the Lewis Structure of a Molecule, Section 1.1, p. 26 Investigation 1.A: Modelling Molecules, Section 1.1, p. 30	Questions for Comprehension: 1-5, Section 1.1, p. 19 Questions for Comprehension: 6, Section 1.1, p. 21 Practice Problems: 1-6, Section 1.1, p. 25 Practice Problems: 6-16, Section 1.1, p. 27 Section 1.1 Review: 3-6, 15, p. 35 Chapter 1 Review: 3-8, 27, 28, 32, p. 44-45 Chapter 2 Test Unit 1 Review: 12, 14, 55, pp. 86-89
20–A2.5k apply VSEPR theory to predict molecular shapes for linear, angular (bent, V-shaped), tetrahedral, trigonal pyramidal and trigonal planar molecules	VSEPR and the Structure of Molecular Compounds, Section 2.1, p. 52–54 Using VSEPR Theory to Predict Molecular Shapes, Section 2.1, p. 55–59 Sample Problem: Predicting Molecular Shape for a Simple Compound, Section 2.1, p. 56	Practice Problems: 1–10, 11–16, Section 2.1, p. 56, 59 Questions for Comprehension: 3, 4, Section 2.1, p. 57 Section 2.1 Review: 5-17, p. 62 Chapter 2 Review: 1, 10, 19, p. 82-83 Chapter Test 2 Test Unit 1 Review: 6, 9, 18, 19-21, 58, 60, 61, pp. 86–89

	Student Textbook	Assessment Options
20–A2.6k illustrate, by drawing or building models, the structure of simple molecular substances	Investigation 1.A: Modelling Molecules, Section 1.1, p. 30 Investigation 2.B: Soap Bubble Molecules, Section 2.1, p.56 Investigation 2.C: Dipole Balloons, Section 2.2, p. 64	Investigation 1.A: Modelling Molecules, Procedure 5, Section 1.1, p. 30 Investigation 2.B: Soap Bubble Molecules: 1-3, Section 2.1, p. 56 Investigation 2.C: Dipole Balloons: 1–4, Section 2.2, p. 64 Chapter 2 Test Unit 1 Review: 6, 9, 18, 19-21, 58, 60, 61, pp. 86–89
20–A2.7k explain intermolecular forces, London (dispersion) forces, dipole-dipole-forces and hydrogen bonding	Investigation 1.A: Modelling Molecules, Section 1.1, p. 30 Dipole-Dipole Forces, Section 2.2, p. 63 Hydrogen Bonding, Section 2.2, p. 64 London (Dispersion) Forces, Section 2.2, p. 67	Investigation 1.A: Modelling Molecules, Analysis 2, Section 1.1, p. 30 Questions for Comprehension: 5–7, Section 2.2, p. 69 Section 2.2 Review: 1–8, p. 70 Chapter 2 Review: 2–8, 11, p. 82 Chapter 2 Test Unit 1 Review: 9, 17, 19, 24, 26, 28, 31–35, 38, 40, 49, 50, 54, pp. 86–89
20–A2.8k relate properties of substances (e.g., *melting and boiling points, enthalpies of fusion and vaporization*) to the predicted intermolecular bonding in the substance	Investigation 2.D: Investigating the Properties of Water, Section 2.2, p. 66 Relating Structures and Properties, Section 2.3, p. 71 Mechanical Properties of Solids, Section 2.3, p. 75 Investigation 2.E: Properties of Substances, Section 2.3, p. 78 Melting Points and Boiling Points, Section 2.3, p. 72 Conductivity, Section 2.3, p. 76	Investigation 2.D: Investigating the Properties of Water: 1–7, Section 2.2, p. 66 Questions for Comprehension: 5–7, Section 2.2, p. 69 Section 2.2 Review: 3, 6-8, p. 70 Questions for Comprehension: 8–10, Section 2.3, p. 75 Section 2.3 Review: 1–10, p. 80 Chapter 2 Review: 2–10, 13, 14, 17–19, 24, 41, p. 82–83 Chapter 2 Test Unit 1 Review: 5, 8-11, 19, 20, 25, 29, 31, 35, 36, 42, 47, 48, 51, 53, 62–65, pp. 86–89
20–A2.9k determine the polarity of a molecule based on simple structural shapes and unequal charge distribution	VSEPR and the Structure of Molecular Compounds, Section 2.1, p. 52 Using VSEPR Theory to Predict Molecular Shapes, Section 2.1, p. 55 Polar Bonds and Polar Molecules, Section 2.1, p. 57	Chapter 1 Review: 22, 23, 24, 35, p. 44-45 Section 2.1 Practice Problems: 1-10, 11-16, p. 56, 59 Questions for Comprehension: 3, Section 2.1, p. 57 Section 2.1 Review: 10-17, p. 62 Chapter 2 Review: 1, 5, 6, 12, 16, 19, 23, p. 82-83 Chapter 2 Test Unit 1 Review: 5, 7, 8 19, 20, 40, 41, 43, 44, 46, 49, 59–61, pp. 86–89
20–A2.10k describe bonding as a continuum ranging from complete electron transfer to equal sharing of electrons.	Bond Type and Electronegativity, Section 1.2, p. 39	Section 1.2 Review: 6, p. 42 Chapter 1 Review: 25, p. 44-45 Chapter 1 Test Unit 1 Review: 4, 9, 20, 40, pp. 86–89

	Student Textbook	Assessment Options
Outcomes for Science, Technology and Society (Emphasis on the nature of science)		
20–A2.1sts explain that scientific knowledge and theories develop through hypotheses, the collection of evidence through experimentation and the ability to provide explanations by ■ *relating chemical properties to their predicted intermolecular bonding by investigating melting and boiling points*	Melting Points and Boiling Points, Section 2.2, p. 72 Mechanical Properties of Solids, Section 2.3, p. 75 Investigation 2.E: Properties of Substances, Section 2.3, p. 78	Chapter 2 Review: 14, 15-18, 28-30, p. 82-83 Unit 1 Review: 45, 62–65, pp. 86–89
20–A2.2sts explain that scientific knowledge is subject to change as new evidence comes to light and as laws and theories are tested and subsequently restructured, revised or reinforced by ■ *explaining how scientific research and technology interact in the production and distribution of beneficial materials, e.g., polymers, household products, life processes, solvents* ■ i*nvestigating how basic knowledge about the structure of matter is challenged in nanotechnology research and development.*	Chemistry File: FYI (sidebar), Section 2.1, p. 60 Network Solids, Section 2.1, p. 60 Career Focus: Ask a Nanotechnologist, p. 84-85	Chapter 2 Review: 14, 15-18, 28-30, p. 82-83 Unit 1 Review: 62–65, pp. 86–89
Skill Outcomes (Focus on scientific inquiry)		
Initiating and Planning		
20–A2.1s ask questions about observed relationships and plan investigations of questions, ideas, problems and issues by ■ stating a hypothesis and making a prediction about the properties of molecular substances based on attractive forces; *e.g., melting or boiling point, enthalpy of fusion, enthalpy of vaporization* ■ describing procedures for safe handling, storage and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labeling information.	Relating Structures and Properties, Section 2.3, pp. 71-74 Investigation 2.D: Investigating the Properties of Water, Section 2.2, p. 66 Investigation 2.E: Properties of Substances, Section 2.3, p. 78	Investigation 2.D: Investigating the Properties of Water, Analysis: 1–6, Conclusion: 7, Section 2.2, p. 66 Investigation 2.E: Properties of Substances, Analysis: 1, Conclusion: 2, Section 2.3, p. 78 Unit 1 Review: 45, 62–65, pp. 86–89
Performing and Recording		
20–A2.2s conduct investigations into relationships among observable variables, and use a broad range of tools and techniques to gather and record data and information by ■ building models depicting the structure of simple covalent molecules, including selected organic compounds ■ carrying out an investigation to determine the melting or boiling point of a molecular substance ■ *using a thermometer and a conductivity apparatus to collect data* ■ *carrying out an investigation to compare the properties of molecular compounds.*	Investigation 2.B: Soap Bubble Molecules, Section 2.1, p.56 Investigation 2.E: Properties of Substances, Section 2.3, p. 78	Investigation 2.B: Soap Bubble Molecules, Analysis: 1, 2, Conclusion: 3, Section 2.1, p.56 Investigation 2.E: Properties of Substances, Analysis: 1, Conclusion: 2, Section 2.3, p. 78 Unit 1 Review: 45, pp. 86–89

	Student Textbook	Assessment Options
Analyzing and Interpreting		
20–A2.3s analyze data and apply mathematical and conceptual models to develop and assess possible solutions by ■ graphing and analyzing data, for trends and patterns, on the melting and boiling points of a related series of molecular substances.	Investigation 2.E: Properties of Substances, Section 2.3, p. 78	Investigation 2.E: Properties of Substances, Analysis: 1, Conclusion: 2, Section 2.3, p. 78 Chapter 2 Review: 21, 27, p. 82-83 Unit 1 Review: 8, 52, 64, pp. 86–89
Communication and Teamwork		
20–A2.4s work as members of a team in addressing problems, and apply the skills and conventions of science in communicating information and ideas and in assessing results by ■ objectively evaluating and analyzing models and graphs constructed by others	Investigation 2.A: Building Ionic Crystals, Section 2.1, p. 51 Investigation 2.B: Soap Bubble Molecules, Section 2.1, p.56	Investigation 2.A: Building Ionic Crystals, Analysis: 2, Section 2.1, p. 51 Investigation 2.B: Soap Bubble Molecules, Analysis: 1, 2, Conclusion: 3, Section 2.1, p.56
■ *researching the ways scientists develop and analyze new materials.*	Investigation 2.E: Properties of Substances, Analysis: 1, Conclusion: 2, Section 2.3, p. 78	Investigation 2.E: Properties of Substances, Analysis: 1, Conclusion: 2, Section 2.3, p. 78

Chapter 2

Diversity of Matter

Student Textbook pages 46–85

Chapter Concepts

Section 2.1 Three-Dimensional Structures

- The shape of the smallest repeating unit in a crystal determines the shape of the entire crystal.
- The Valence-Shell Electron-Pair Repulsion (VSEPR) theory allows you to predict the three-dimensional structure of simple molecules.

Section 2.2 Intermolecular Forces

- Some molecules form permanent dipoles, where the positive end of one molecule is attracted to the negative end of another.
- Hydrogen bonds can occur between molecules when, in each molecule, a hydrogen atom is covalently bonded to an atom with a high electronegativity.
- All molecules are attracted to one another by London (dispersion) forces. In non-polar molecules, these forces are the only attractive forces.

Section 2.3 Relating Structures and Properties

- The state of a substance depends on the strength of the forces between molecules and the average kinetic energy (temperature) of the particles.
- Melting and boiling points of pure substances are an indication of the nature of the bonding in the substance.
- The conductivity of a substance depends on the freedom of electrons or ions to move past one another.

Common Misconceptions

- The formula of an ionic compound such as NaCl is referred to as a *formula unit*, not a molecule. The term *molecule* refers to covalently bonded compounds only. Insist that students distinguish between these terms and use them correctly.
- Students may believe that covalent compounds that contain polar covalent bonds always result in polar molecules. Use the example of carbon dioxide to stress that this is not necessarily the case. Students may also believe that molecules that are shaped like other polar molecules (water or ammonia, for example) must be polar molecules. This is only the case when they contain polar covalent bonds.
- Because both involve charges, students may confuse the attraction between polar covalent molecules with the attraction between ions in an ionic compound. Re-emphasize the difference between intermolecular bonds and intramolecular bonds, and ensure students realize that the charges across polar covalent bonds are only partial charges.
- Students may assume that, since ionic compounds tend to have high melting and boiling points and covalent compounds tend to have low melting and boiling points, ionic bonds are a great deal stronger than covalent bonds. Ensure students understand that, in a solid ionic compound, strong ionic bonds are present and extend throughout the crystal lattice. There is no difference between intermolecular forces and intramolecular bonds in an ionic crystal. This is also true for a network solid such as diamond. In the case of solid covalent compounds however, strong covalent bonds hold atoms together as molecules, while relatively weak intermolecular forces hold the molecules to one another as a solid.
- Students can easily confuse the role of intermolecular forces with intramolecular bonding. Physical changes involve the re-arrangement of intermolecular forces, while chemical changes involve the re-arrangement of intramolecular bonds. For example, it is the intermolecular bonds that are broken as a solid covalent compound melts or vaporizes. Emphasize that when a solid covalent compound melts, individual molecules remain unchanged. For example, when ice melts to form liquid water, the water molecules are unchanged because it is the forces between water molecules that have been broken. If the bonds within water molecules were broken during this physical change, melting water would produce hydrogen and oxygen gases.

Helpful Resources

Books and Journal Articles

- Erduran, Sibel, "Philosophy of Chemistry: An Emerging Field with Implications for Chemistry Education," *Science & Education,* 10, pp. 581–593. (Professor at King's College, University of London, includes an in-depth study of the uses and pitfalls of models)
- Gillespie, Ronald J., "Teaching Molecular Geometry with the VSEPR Model," *Journal of Chemistry Education,* 2004, Vol. 81, No. 3, p. 298.
- Harrison, Allan G. and Treagust, David F., "Learning About Atoms, Molecules, and Chemical Bonds: A Case Study of Multiple-model Use in Grade 11 Chemistry," *Science Education,* Vol. 84, Issue #, pp. 352–381
- Peterson, R.F., and Treagust, D., "Grade 12 Students' Misconceptions of Covalent Bonding and Structure," *Journal of Chemistry Education,* 1989, Vol. 66. No. 6, p. 459.
- Pittman, Faith A., "Intermolecular Forces as a Key to Understanding the Environmental Fate of Organic Xenobiotics," *Journal of Chemistry Education,* 2005, Vol. 82, No.2, p. 260.

Web Sites

Web links related to diversity of matter can be found at **www.albertachemistry.ca**. Go to the Online Learning Centre, and log on to the Instructor Edition. Choose Teacher Web Links.

List of BLMs

Blackline masters (BLMs) have been prepared to support the material in this chapter. The BLMs are either for assessment (AST); use as overheads (OH); use as handouts (HAND), in particular to support activities; or to supply answers (ANS) for assessment or handouts. The BLMs are in digital form, stored on the CD-ROM that accompanies this Teacher's Resource or on the web site at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs. The BLMs can be modified to suit the needs of your students.

Number (Type) Title
2.0.1 (HAND) Launch Lab: Crystalline Columns
2.0.1A (ANS) Launch Lab: Crystalline Columns Answer Key

2.1.1 (OH) 3-D Structure of Crystals
2.1.2 (OH) Sodium Chloride Crystals
2.1.3 (OH) Comparison of Unit Cells
2.1.4 (HAND) Investigation 2.A: Building Ionic Crystals
2.1.4A (ANS) Investigation 2.A: Building Ionic Crystals Answer Key
2.1.5 (OH) VSEPR Shapes: Linear
2.1.6 (OH) VSEPR Shapes: Trigonal Planar
2.1.7 (OH) VSEPR Shapes: Bent
2.1.8 (OH) VSEPR Shapes: Tetrahedral
2.1.9 (OH) Molecular Shapes
2.1.10 (HAND) Investigation 2.B: Soap Bubble Molecules
2.1.10A (ANS) Investigation 2.B: Soap Bubble Molecules Answer Key
2.1.11 (OH) Dipoles and Polarity
2.1.12 (OH) Molecular Shapes and Polarity

2.2.1 (OH) Intermolecular Forces: Dipole-Dipole Attractions
2.2.2 (HAND) Investigation 2.C: Dipole Balloons
2.2.2A (ANS) Investigation 2.C: Dipole Balloons Answer Key
2.2.3 (OH) Intermolecular Forces: Hydrogen Bonding
2.2.4 (HAND) Investigation 2.D: Investigating the Properties of Water
2.2.4A (ANS) Investigation 2.D: Investigating the Properties of Water Answer Key
2.2.5 (OH) Induced Dipoles
2.2.6 (OH) Comparing Dispersion Forces in Isotopes

2.3.1 (OH) Bonding in Solid, Liquid, and Gas Phases
2.3.2 (OH) Melting and Boiling Points of Some Common Compounds
2.3.3 (OH) Boiling Points of Binary Hydrides
2.3.4 (OH) Conductivity of Ionic Compounds
2.3.5 (HAND) Investigation 2.E: Properties of Substances
2.3.5A (ANS) Investigation 2.E: Properties of Substances Answer Key
2.4.1 (AST) Chapter 2 Test
2.4.1A (ANS) Chapter 2 Test Answer Key

Using the Chapter 2 Opener

Student Textbook pages 46–47

Teaching Strategies

- To help students visualize the nature of ionic and molecular bonds, have them play the roles of the nuclei and electrons in some simple compounds. Dramatically demonstrating gaining, losing, and sharing electrons in this way will help fix the concepts in their minds and will support further discussions of lone electron pairs, bonding pairs, and co-ordinate covalent bonding. You might even give a simple introduction to VSEPR theory in this way by having students demonstrate what happens when the electrons in the bond repel each other.

Launch Lab Crystalline Columns

Student Textbook page 47

Purpose

Using a burette and stand, students will slowly drip a solution of sodium acetate trihydrate and water onto a sodium acetate trihydrate crystal to observe the creation of crystals similar to a stalagmite in a cave. They will answer questions about the formation and shapes of crystals.

Outcomes

20–A1.6k

Advance Preparation

When to Begin	What to Do
2–3 weeks before	■ Ensure availability of sodium acetate trihydrate crystals and necessary equipment, including balance, graduated cylinders, flask, hot plates, water bottles, burettes and stands, glass plates, and forceps.
1 day before	■ Collect apparatus. ■ Set up burettes and stands with a glass plate for each. ■ Photocopy **BLM 2.0.1 Launch Lab: Crystalline Columns.**

Materials

- balance
- 10 mL graduated cylinder
- 100 mL Erlenmeyer flask
- hot plate
- water bottle
- burette and stand
- glass plate
- forceps
- water
- sodium acetate trihydrate crystals

Time Required

60 minutes

Helpful Tips

- Sodium acetate trihydrate is often simply called sodium acetate. Its formula is $CH_3COONa•3H_2O$. In solution, it reacts with water to form acetic acid.
- Ensure that students heat the mixture only until the crystals are dissolved. The result will be a saturated solution that, as it cools, will become supersaturated, resulting in the formation of crystals. The melting point of sodium acetate trihydrate is 58 °C. Its solubility is excellent. Prolonged heating may result in production of harmful fumes.
- Advise students that all crystals must be dissolved before proceeding with the next step. Any undissolved crystal will act as a seed to start a crystallization process, which is the purpose of the crystal placed on the plate.
- If this is your first time with this procedure, perform the procedure yourself ahead of time so that you know what to expect, how high to raise the burette safely, how long to heat the mixture, how fast to drip out the solution, and so on.
- Use **BLM 2.0.1 (HAND) Launch Lab: Crystalline Columns** to support this activity. Remove sections as appropriate to meet the needs of the students in your class. The answers to questions are on **BLM 2.0.1A (ANS) Launch Lab: Crystalline Columns Answer Key**.
- ***Expected Results*** The sodium acetate trihydrate crystal on the glass plate acts as a seed. As the nearly saturated solution drops from the burette onto the crystal and the water evaporates slightly, more sodium acetate will crystallize onto the seed crystal. A larger, possibly columnar crystal should grow. As well, evaporation of water from the tip of the burette might allow crystals to begin to form on the burette. A long crystal might grow down from the burette.

Safety Precautions

- Ensure students read and understand “Safety in Your Chemistry Laboratory and Classroom” in the student textbook.
- Point out the safety precautions icons and ensure that students avoid fumes and wear safety glasses, aprons, and gloves to protect clothing and skin.
- The procedure directs that the burette be raised as high as it will safely go. Ensure that the stand and burette are stable.
- Sodium acetate trihydrate crystals may be harmful if ingested, inhaled, or absorbed through the skin. Symptoms in extreme cases may include coughing, sore throat, laboured breathing, chest pain, and irritation with redness and pain to skin and eyes. If necessary, remove the person to fresh air, or in the case of contact to eyes or skin, flush with plenty of water.
- Dispose of all materials safely. Unused solution may be disposed of in the sink since it is similar to vinegar when diluted.

Answers to Analysis Questions

1. Students should observe the formation of a solid on the glass plate. The mass of the solid grows larger with each drop of solution.
2. Sodium acetate trihydrate is ionic.
3. As the solution was dripped onto the crystal, more ions were added to the ions already bonded together, increasing the size of the compound mass. Water evaporated from the drops as they fell, causing the ions to come out of solution readily.
4. The crystal on the plate provides beginning ions for the ions in the solution to bond to. It is a “seed crystal.”
5. Stalagmites and stalactites are formed from drip-water in caves that contain high concentrations of mineral ions, (mostly calcium carbonate). As the water evaporates, the ions precipitate out. Dripping causes successive layers to form as the ions of $CaCO_3$ bond with the crystals already formed, adding new layers.
6. Possible answers might include: the size of the atoms or ions will be different in different compounds; the bonds will vary according to the ions and electronegativities. Students might suggest that the composition of the atoms and differing bonds will result in different shapes for the smallest unit of the compound.

Assessment Options

- Assess students’ participation in the class discussion about this activity.
- Have students write what they already know about ionic compounds, what they learned about the formation of crystals of ionic compounds, and questions they have that were triggered by the activity. Assess based on completeness and thoughtfulness (not correctness) of responses.
- Collect and assess answers to Analysis Questions.

2.1 Three-Dimensional Structures

Student Textbook pages 48–62

Section Outcomes

Students will:

- build models of common ionic lattices
- analyze models of ionic compounds built by others
- apply VSEPR theory to predict the shapes of molecular compounds
- determine the polarity of a molecule based on simple structural shapes and unequal charge distribution

Key Terms

formula unit
Valence-Shell Electron-Pair Repulsion (VSEPR) theory
linear
bond angle
trigonal planar
tetrahedral
pyramidal
bent
network solids

Chemistry Background

- Crystals are made up of repeated patterns of one of 14 different kinds of unit cells. There are three types of cubic unit cells: simple cubic, body-centred cubic, and face-centred cubic. The sodium chloride unit cell is simple cubic.
- Crystals themselves are classified into seven groups or systems of unit cells, based on the relative lengths of the sides of their unit cells and the angles between the faces. The names of the systems are cubic, tetragonal, orthorhombic, monoclinic, triclinic, hexagonal, and rhombohedral.
- VSEPR theory is based on the idea that like charges repel each other. The negatively charged electrons in any bonding pair (BP) of electrons will repel and be repelled by the other negatively charged pairs of electrons in other bonding pairs and/or lone pairs (LPs) around the atom. The molecular shape allows the pairs to be as far apart as possible, while still being held by the attractive forces of the atoms' positive nuclei.
- The student textbook gives only an introduction to VSEPR theory. The theory is much more complex than outlined here, but this treatment will give students a feel for using VSEPR to determine molecular shapes. Other shapes not described include trigonal bipyramidal, octahedral, square planar, and T-shaped.
- The formation of a network solid when silicon and oxygen combine, instead of the SiO_2 molecule that an electron dot diagram might predict, shows the necessity of experimentally confirming theoretical predictions. Several geometric isomers and/or allotropic forms of substances are possible.

Teaching Strategies

- Recently, students will have finished the previous chapter, which was mostly a review. For security, they may wish to cling to the things they know well from *Science 9* and *Science 10*. Challenge them to use that base and expand their understanding by anticipating, stepping forward, and accepting the extensions of that knowledge seed into a growing and vibrant plant.
- Emphasize that structural models of ionic bonds and molecular compounds are just models. Point out the characteristics of different types of models and compare them to observed properties. The most effective approach in terms of cognitive understanding of bonding and structures has proven to be a multi-model approach.
- **www.chempensoftware.com** has a 30-day trial version of a molecule-building program called ChemPen3DTM, which analyzes bond angles and bond length, and creates illustrations for almost every bond combination on the periodic table, including metals to metals.
- A number of overhead masters have been prepared for this section. You will find them with the Chapter 2 BLMs on the CD-ROM that accompanies this Teacher's Resource or at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs.

  Number (Type) Title
  2.1.1 (OH) 3-D Structure of Crystals
  2.1.2 (OH) Sodium Chloride Crystals
  2.1.3 (OH) Comparison of Unit Cells
  2.1.5 (OH) VSEPR Shapes: Linear
  2.1.6 (OH) VSEPR Shapes: Trigonal Planar
  2.1.7 (OH) VSEPR Shapes: Bent
  2.1.8 (OH) VSEPR Shapes: Tetrahedral
  2.1.9 (OH) Molecular Shapes
  2.1.11 (OH) Dipoles and Polarity
  2.1.12 (OH) Molecular Shapes and Polarity

Answers to Questions for Comprehension

Student Textbook page 50

Q1. Two characteristics of ions that determine how they pack together to form crystals are the ion charges and their relative sizes. How quickly the crystal forms will also affect packing.

Q2. The main reason why ions pack together with different types of symmetry is the relative charge of the ions. In the sodium chloride crystal, both sodium and chloride ions each have a charge of 1. So, they will pack in a ratio of 1:1. If the crystal is made of calcium and chloride ions, however, the ratio must be 1:2, since the charge on a calcium ion is 2^+, and the charge on chloride ion is only 1^-. The difference in the ratio will have a major effect on the crystal shape.

Investigation 2.A: Building Ionic Crystals

Student Textbook page 51

Purpose

Students will build models of common ionic lattices to represent real crystals. They will work as teams and critically analyze the models they have created. They will also critically analyze the models built by their classmates.

Outcomes

- 20-A1.2s
- 20-A1.4s

Advance Preparation

When to Begin	What to Do
2–3 weeks before	■ Ensure the availability of the polystyrene balls or gumdrops.
1 day before	■ Collect materials. ■ Divide materials into quantities appropriate for the number of groups in your class. ■ Photocopy **BLM 2.1.4 (HAND) Investigation 2.A: Building Ionic Crystals.**

Materials
■ polystyrene (Styrofoam™) balls of different sizes or gumdrops in two different sizes or colours ■ toothpicks

Time Required

35–40 minutes

Helpful Tips

- Emphasize with students that models do not represent molecules, but illustrate the bonding of ions to ions in the formula ratios. Remind students that the toothpicks do not represent actual connections between ions.
- If you use gumdrops, ensure that the students do not eat the gumdrops distributed for making models, either before or after they construct their models. Repeated handling will make them unsanitary for eating. Arrange to have extras for students to eat after the investigation is completed.
- Use **BLM 2.1.4 (HAND) Investigation 2.A: Building Ionic Crystals** to support this activity. Remove sections as appropriate to meet needs of students in your class.

SUPPORTING DIVERSE STUDENT NEEDS

Assign partners to pair up stronger students with weaker students for peer coaching.

Answers to Analysis Questions

1. Answers may include converting a two-dimensional illustration to a three-dimensional model, and visualizing ions, bonds, or the correct angles to join the "ions" together.
2. You would not expect a wide variation in the models produced by the students. If there are apparent differences, they should be caused by the difference in angles. The ion-to-ion ratios should be the same in every case.
3. By building three-dimensional models of crystals, students are able to see all of the faces and edges of the crystal by rotating the model in space. This appreciation is not obtained by looking only at two-dimensional pictures.

Answer to Conclusion Question

4. The following are some of the ways that models represent real crystals: they show individual ions in the same ratio; they illustrate the electrostatic attractions between ions; and they show the angles between the ions. The following are some of the ways that models are different from real crystals: atoms are mostly empty space; ions are not hard; bonds are not solid entities but represent an attraction between the nuclei of ions and electrons; and relative sizes are not always accurately represented.

Assessment Options

- Assess students' participation in the class discussion surrounding this investigation.
- Use Assessment Checklist 5: Learning Skills to assess how students applied previous knowledge to the project and how they performed with others.
- Use Assessment Checklist 4: Performance Task Group Assessment, to help students assess how they worked together as a group.
- Collect student responses to Conclusion Question 4. Assess for completeness, thoughtfulness, and understanding of the structures of ionic crystals.

Answers to Practice Problems 1–10

Student Textbook page 56

For full solutions to the practice problems, visit **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, Full Solutions.

1. linear
2. tetrahedral (pyramidal)
3. trigonal planar

4. trigonal planar
5. tetrahedral (pyramidal)
6. tetrahedral (bent)
7. tetrahedral
8. tetrahedral
9. linear
10. trigonal planar

Investigation 2.B: Soap Bubble Molecules

Student Textbook pages 56–57

Purpose

Using different kinds of models of molecules is most effective in developing cognitive understanding of molecular structure. Soap bubbles produce good models of molecular shapes. Students will blow soap solution through straws to produce mounds of soap bubbles of different shapes and sizes to simulate bonding pairs and lone pairs of electrons. They will also examine the simulated bond angles.

Outcomes

- 20–A2.4k
- 20–A2.4s

Advance Preparation

When to Begin	What to Do
1 month before	■ Ensure materials are available. ■ Prepare soap solution.
1 day before	■ Discuss procedure with students. ■ Have students write up procedure. ■ Photocopy **BLM 2.1.10 (HAND) Investigation 2.B: Soap Bubble Molecules.**

Materials
■ 100 mL beaker ■ hard, flat surface ■ straw ■ two transparent 15 cm rulers ■ protractor ■ paper towels ■ soap solution (mixture of proportions: 80 mL distilled water, 15 mL dish soap, and 5 mL glycerin), enough for each group to have 25 mL

Time Required

40 minutes

Helpful Tips

- To prepare 3.0 L of soap solution, put 2.4 L of distilled water in a clean container you can seal. Add 450 mL liquid detergent and 150 mL glycerin. Do not shake or agitate. Let the soap stand for a month.
- Practise making the soap structures ahead of time so that you know what size of bubbles work best, how hard to blow, and so on. Demonstrate the method to the students before they begin.
- Assign roles such as “procedure navigator” to make sure all the steps are completed, “angle measurer” to record bond angles, “recorder,” and so on.
- Clarify that Procedure step 4 might be done in two different ways: linear, as in CO_2, or trigonal planar, as in BF_3. In order to go on to step 5, students will have to use the trigonal planar shape.
- Use **BLM 2.1.10 (HAND) Investigation 2.B: Soap Bubble Molecules** to support this activity. Remove sections as appropriate to meet the needs of the students in your class. The answers to questions are on **BLM 2.1.10A (ANS) Investigation 2.B: Soap Bubble Molecules Answer Key.**
- ***Expected Results*** The surface tension of the soap bubble film will tend to create the smallest surface to volume ratio possible. This simulates the shape of “space filling” atoms in a molecule. When there are three bubbles about the same size, the angles between bubbles should be 120°. When the third bubble is larger than the first two, the angle between the planes made by the larger bubble and the two smaller bubbles will be larger than 120° and the angles made by the planes of the smaller bubbles will be less than 120°. It will be a little more difficult to measure the angles formed by four bubbles but in every case, the planes on the larger bubbles will form larger angles.

Safety Precautions

- Ensure that students follow the safety precautions in the student textbook.
- Each person should use a clean straw and avoid getting soap solution in the mouth.
- Wipe up any spills immediately.

Answers to Analysis Questions

1. The shape associated with two same-sized bubbles is linear, and the bond angle is 180°. Examples of molecules with this shape include: CO_2 and $BeCl_2$. The shape associated with three same-sized bubbles is trigonal planar, and the bond angle is 120°. An example of a molecule with this shape is BCl_3. The shape associated with four same-sized bubbles is tetrahedral, and the bond angle is 109.5°. An example of a molecule with this shape is CCl_4.

2. In step 6, there are two small bubbles and one slightly larger bubble. An example is $SnCl_2$. In step 7 there are two same-sized bubbles and two slightly larger bubbles. An example of this shape is NH_3.

Answer to Conclusion Question

3. Students may say that the soap molecules tend to be attracted to each other by trying to minimize surface area while maintaining their volume. This causes them to adopt their respective positions and shapes. A more scientific explanation is that a soap bubble has both polar portions (carboxyl groups) and non-polar portions (hydrocarbon portions). The non-polar ends are attracted to each other, tending to reduce the size of the bubble. The polar ends of a soap bubble point outward, away from the inside of the bubble, and act in a similar way to electrons in an atom.

Assessment Options

- Assess students' participation in the discussion surrounding this activity.
- Use Checklist 3: Performance Task Self-Assessment and Checklist 4: Performance Task Group Assessment to assess students' participation in their groups (See Appendix A.)
- Collect and assess students' write-ups using Checklist 2: Laboratory Report.

Answers to Questions for Comprehension

Student Textbook page 57

Q3. The features that must be obtained from the Lewis structures are the types of bonding pairs (LP or BP, how many of each), and whether the bonds are single, double, or triple.

Q4. "Electron group" is defined as the electrons involved in a bond. It may be a single, double, or triple bond.

Answers to Practice Problems 11–16

Student Textbook page 59

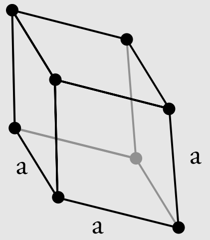
For full solutions to the practice problems, visit **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, Full Solutions.

11. tetrahedral, polar
12. trigonal planar, polar
13. tetrahedral (pyramidal), polar
14. research shows that depending on the bends, the vectors may or may not cancel
15. tetrahedral, polar
16. $NF_3(g)$ more polar than $NCl_3(\ell)$

Section 2.1 Review Answers

Student Textbook page 62

1. A unit cell is defined as the smallest set of ions in a crystal for which a pattern is repeated over and over. It is a three-dimensional structure that details how each of the faces will appear in the crystal. Sodium chloride, NaCl, has a cube-shaped unit cell (as shown in Figure 2.2(C) on page 48), whereas calcium carbonate, $CaCO_3$, has a rhombohedral shape (diamond), as shown below:



2. Cations (positive ions) are smaller than neutral atoms of the same element, whereas anions (negative ions) are larger than neutral atoms of the same element. The cations in any period of the periodic table are smaller than the anions of the same period.
3. The two factors that affect the shape of an ionic crystal are the relative size and charge of the ions in the crystal.
4. In an ionic crystal, the charged ions are simultaneously attracted to all of the oppositely charged ions surrounding them, which leads to the building of a crystal of potentially infinite size. In a molecule, however, the attractions, or bonds that hold the atoms in the molecule together, exist only between specific pairs of atoms.
5. VSEPR theory states that bonds and lone pairs spread out as much as possible in a molecule. By counting the number and type of bonding elements, the shape of the molecule can be predicted because it will maximize the distance between the bonding elements.
6. LP-LP (lone pair-lone pair) > LP-BP (bonding pair) > BP-BP. The order of repulsion will not affect the overall shape of the molecule but will affect the bond angle between the atoms in the molecule. For example, the theoretical angle between any two atoms in a tetrahedral shape is 109.5º. In a bent molecule, the bond angle would be 109.5º if lone pair repulsions were not greater than either lone pair-bonding pair or bonding pair-bonding pair repulsions. However, in the bent molecule (water (H_2O)) the bond angle is 104.5º because of the greater repulsions between the lone pairs and the lone pair-bonding pairs than between the bonding pairs.
7. The type of bond, whether single, double, or triple does affect the molecular shape. A triple bond, although consisting of three bonding pairs, is only one

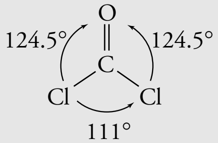
bonding element. One bonding element leads to a linear shape, whereas three single bonds, or three bonding elements, would be either pyramidal (with one LP) or trigonal planar (no LP).

8.

Shape	Configuration leading to the shape	Types of bonds
Tetrahedral	4 bonding pairs, no lone pairs	4 single bonds
Trigonal Planar	3 bonding pairs 4 bonding pairs	3 single bonds, no lone pairs 1 double bond, 2 single bonds
Pyramidal	3 bonding pairs, one lone pair	3 single bonds
Bent	2 bonding pairs, 2 lone pairs 3 bonding pairs, 1 lone pair	2 single bonds 1 double bond, 1 single bond
Linear	1 bonding pair, 3 lone pairs 2 bonding pairs, 2 lone pairs 3 bonding pairs, 1 lone pair	1 single bond 1 double bond 1 triple bond

9. In a trigonal planar molecule, there are three electron groupings which are all bonding pairs. The ideal angle between the bonds is 120°.

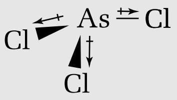
(a)



(b) The predicted angles of 120° are based on each of the repulsions being equal. Based on the observed bond angles, we could infer that the repulsions from a double bond are greater than the repulsions from a single bond, leading to the deviations observed.

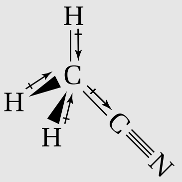
(c) The bond angles between the oxygen and hydrogen in methanal are 121°, compared to 124.5° observed in this molecule. The repulsion between the oxygen and chlorine atoms is greater than the repulsion between oxygen and hydrogen atoms. Since chlorine has three lone pairs, there should be more repulsion between oxygen and chlorine, as observed.

10. (a)



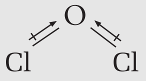
Trigonal pyramidal molecule: the top (arsenic end) of the pyramidal shape is positive, the base vertices (chlorine ends) are negative. In each As—Cl bond, the As is δ^+ and the Cl is δ^-.” The lone electron pair at the top of the pyramidal shape repels the bonding pairs. The molecule is polar.

(b)



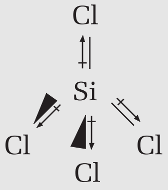
Tetrahedral molecule: The H ends of the C—H bonds are δ^+, giving those ends partial positive charge. The nitrogen end of the C—N triple bond is δ^- and tends to have a partial negative charge, making the molecule polar.

(c)



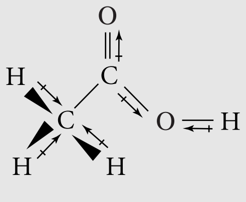
The central oxygen atom has two LPs and two BPs. The two LPs repel each other and the chlorine BPs, so the shape is bent. The chlorine ends of the bonds are δ^+ and the oxygen end is δ^-, making the molecule polar.

(d)



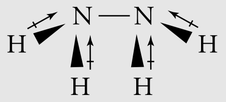
Even though each Si—Cl bond is polar, the individual Cl atoms surrounding the Si central atom exert an identical force on the bonding electrons. The Si—Cl bonds are symmetrically arranged, so the shape is tetrahedral. The polarity of the individual Si—Cl bonds neutralize each other to give a zero net polarity, so the molecule is non-polar.

(e)



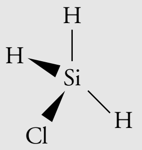
The molecule is tetrahedral around the first carbon and trigonal planar around the second carbon. The molecule is polar, with the oxygen COOH end of the molecule δ^- and the CH_3 end δ^+.

(f)



The molecule is trigonal pyramidal around each of the two nitrogen atoms. Although the formula makes it look symmetrical, the shape indicates that it is not. The molecule is polar.

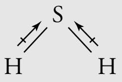
(g)



Shape: tetrahedral

Polarity: polar

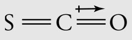
(h)



Shape: bent

Polarity: polar

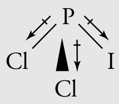
(i)



Shape: linear

Polarity: polar

(j)



Shape: pyramidal

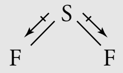
Polarity: polar

11. This molecule is based on a tetrahedral. Although the bonds themselves are slightly polar (since the difference in electronegativities between hydrogen and carbon is 0.4), the molecule itself is non-polar because it is symmetrical.

12. All polar molecules must have polar bonds, but not all non-polar molecules have non-polar bonds. If there is symmetry in a molecule such that the polarity of the individual bonds cancel each other to give a zero net charge, the molecule is non-polar. For example, CCl_4 is a non-polar molecule that has four polar bonds.

13. The number of bonding pairs and non-bonding pairs of electrons around the central atom of the molecule affects the shape of the molecule. The polarity of the individual bonds and the shape of the molecule affect its polarity.

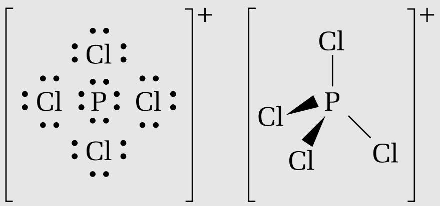
14.



Since $\Delta EN = 3.98 - 2.58 = 1.40$, the S—F bond is polar. The four electron pairs are arranged in a tetrahedral shape, and the shape of the molecule is bent. Thus, the S—F bonds do not cancel each other, and the molecule is polar.

15. Both CH_4 and CH_3OH have a tetrahedral electron pair arrangement and a tetrahedral shape. CH_4 is non-polar because of the symmetry of the molecule, with four hydrogens around the central atom. CH_3OH is polar because the polarity of the CO—OH bond is different from the polarity of the three C—H bonds.

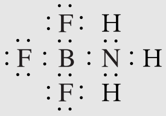
16.



Since there is symmetry in the four P—Cl bonds, the electron pairs are arranged in a tetrahedral shape and the shape of the ion is also tetrahedral.

17. The combination reaction can be represented by the following equation:

$$BF_3 + NH_3 \rightarrow F_3BNH_3$$



Before the reaction, boron only has six electrons in the outermost energy level. The product is expected to be more stable than the reactants, because when nitrogen's lone pair is shared with boron, the covalent B—N bond completes the stable configuration of the nearest noble gas for each element in the compound.

2.2 Intermolecular Forces

Student Textbook pages 63–70

Section Outcomes

Students will:

- explain dipole-dipole attractions, hydrogen bonding, and London (dispersion) forces
- explain how research and technology interact in the production of beneficial materials

Key Terms

intramolecular forces
intermolecular forces
dipoles
dipole-dipole attraction
hydrogen bond
dispersion force
London force

Chemistry Background

- Van der Waal's forces exist *between* molecules, and are termed *inter*molecular forces. They are usually classified as dipole-dipole forces, hydrogen bonds, and London (dispersion) forces. Although there is some overlap in strengths, generally hydrogen bonds are stronger than dipole-dipole forces, and London (dispersion) forces are the weakest.

- The molecular bonds studied up to this point, non-polar covalent and polar covalent, are bonds *within* molecules and are *intra*molecular bonds. Students will have trouble distinguishing between *intra*molecular bonds and *inter*molecular forces, and will need to be reminded which type you are talking about.
- Hydrogen bonds only occur between molecules in which hydrogen is bonded with oxygen, fluorine, or nitrogen. It occurs with these molecules because they have the greatest electronegativities and the bond with hydrogen is highly polar. Hydrogen has no inner electrons to shield the nucleus, and in a highly polar bond, this leaves the single proton largely exposed. Thus, the hydrogen nucleus can be attracted to the lone pairs in another atom. This attraction is only about 5% as strong as a covalent bond, but significant enough that it makes life on Earth possible, due to the unique properties it gives to water.
- There are also electrostatic forces of attraction or repulsion between ions and dipoles, between ions and non-polar molecules, and between dipoles and non-polar molecules.

Teaching Strategies

- Bring samples of salt and sugar to class. Discuss the appearance, taste, molecular or ionic structure, and properties of each, what is similar and what is different. Discuss the bonding in both substances. Emphasize that intramolecular bonding cannot account for all of the properties of the substances.
- As an interesting demonstration of how the strength of bonds affects properties, place one marble in each of four test tubes. Add one of the following to each of the test tubes: water, ethanol, antifreeze (ethylene glycol), glycerin. Seal each test tube with a rubber stopper. Turn all four test tubes upside down at the same time. The rate of fall of the marble through each liquid gives a measure of the relative viscosities, and therefore the strength of the hydrogen bonding in each liquid.
- Overhead masters have been prepared for this section. You will find them with the Chapter 2 BLMs on the CD-ROM that accompanies this Teacher's Resource or at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs.
  Number (Type) Title
  2.2.1 (OH) Intermolecular Forces: Dipole-Dipole Attractions
  2.2.3 (OH) Intermolecular Forces: Hydrogen Bonding
  2.2.5 (OH) Induced Dipoles
  2.2.6 (OH) Comparing Dispersion Forces in Isotopes

Investigation 2.C: Dipole Balloons

Student Textbook page 64

Purpose

Students will use electrostatically charged balloons to model dipoles in molecules. Balloons can be considered dipoles when one end is charged and one end remains uncharged.

Outcomes

- 20–A2.4k
- 20–A2.7k
- 20–A2.1s

Advance Preparation

When to Begin	What to Do
2–3 weeks before	■ Ensure that materials are available.
1 day before	■ Have students read the entire procedure and make predictions about the situations described in the procedure. ■ Assign partners. ■ Photocopy **BLM 2.2.2 (HAND) Investigation 2.C: Dipole Balloons.**

Materials
■ round balloons, 2 per student pair ■ string ■ marker

Time Required

30 minutes

Helpful Tips

- Use **BLM 2.2.2 (HAND) Investigation 2.C: Dipole Balloons** to support this activity. Modify it as necessary.
- The balloons do not work well if either over- or under-inflated. Experiment ahead of time so that you have a good idea about the level of inflation that works best.
- Make sure students understand not only how balloons are models of charged molecules (Conclusion question 4), but how they are different. A polar molecule has a positive and a negative end. When the molecule orients itself toward other molecules, the negative and positive ends rotate in alignment with the negative and positive ends of the other molecules, rather than the whole molecule moving away as the balloons do.
- ***Expected Results*** When you move your hand near the charged spot on one balloon, the balloon will move slightly toward your hand. The reason for the movement is that the charge on the balloon induces a charge on the surface of your skin and the opposite charges attract each other.

When you more the charged spots of two balloons toward each other, they will tend to move apart because the like charges on the balloons will repel each other. (They might twist a bit and move the charged spots so they are not facing each other.) When you place your hand between the two balloons that are repelling each other, they will not repel each other as much. In fact, they might move toward your hand. The explanation for this movement is that the charges on the balloons are inducing dipoles in the molecules on your skin, causing opposite charges (to those on the balloons) to move toward the surface of your skin and the balloons to then be attracted to your skin.

Answers to Analysis Questions

1. When a hand moves close to the negatively charged end of the balloon, any polar molecules in the hand will rotate so that the positive ends of the molecules will be toward the balloon. Also, there will be a net migration of free electrons away from the area closest to the balloon to farther parts of the body. The hand will be left with a net positive charge. This is called charging by induction. These effects are slight but will be enough to attract the negatively charged balloon toward the hand.
2. Since the charged areas of the balloons are both negative, and the charged areas are facing each other, the balloons will move apart because of the repulsion effect of like charges.
3. Your hand still has a net positive charge. Just as it attracted one negatively charged balloon, it will also attract balloons from either side. Both balloons will move toward your hand.

Answer to Conclusion Question

4. The negative ends of the balloons repel and are repelled by each other and are attracted to neutral and positive entities. In this way they are like polar molecules. The balloons are unlike polar molecules because the balloons have a net negative charge, but polar molecules are electrically neutral, because their positive and negative ends cancel each other.

Assessment Options

- Use Assessment Checklists 2, 3, 4, and 5 as assessment aids (see Appendix A.)
- Collect and assess written responses to Analysis and Conclusion questions. Assess for understanding, completeness, clarity, and effort.

Investigation 2.D: Investigating the Properties of Water

Student Textbook pages 66–67

Purpose

Students will explain the properties of molecular substances based on attractive forces. In *Science 10,* students encountered hydrogen bonding in studying the properties of water. However, they are much more able in this investigation to understand hydrogen bonds, and other intermolecular forces based on positive and negative charges, as they relate to the properties of water and to properties of molecular substances.

Outcomes

- 20–A2.9k
- 20–A2.1s

Advance Preparation

When to Begin	What to Do
2–3 weeks before	■ Ensure availability of all materials.
1 day before	■ Collect the materials and apparatus. ■ Run through the procedure yourself so you know what to expect from the students. ■ Photocopy **BLM 2.2.4 (HAND) Investigation 2.D: Investigating the Properties of Water.**

Materials

- 500 mL beaker
- two shallow dishes
- sewing needles
- two small glasses
- 150 mL beaker
- acetate strip
- cotton cloth
- vinyl strip
- wool cloth
- vegetable oil
- water
- ethanol
- pepper
- liquid dish detergent

Time Required

30 minutes

Helpful Tips

- Students have studied this topic before and may say that they have already done an investigation on the properties of water. You may want to remind them of this ahead of time and emphasize that they will be looking more closely at some properties and will be asked to explain what they see in terms of intermolecular forces such as hydrogen bonding and polarity.
- Collect and assess written responses to Analysis and Conclusion questions. Assess for understanding, completeness, clarity, and effort. Use these also as a guide to assess which concepts need clarification.
- Use **BLM 2.2.4 (HAND) Investigation 2.D: Investigating the Properties of Water** to support this activity. Remove sections as appropriate to meet the needs of the students in your class. The answers to questions are on **BLM 2.2.4A (ANS) Investigation 2.D: Investigating the Properties of Water Answer Key.**
- ***Expected Results*** When either the charged acetate strip or vinyl strip are held near the small stream of water, the bends bring the water nearer to the strips. Although the two strips are charged oppositely, the results are the same. The positively charged acetate strip attracts the negative poles of the water molecules and the negatively charged vinyl strip attracts the positive poles of the water. Because oil molecules are not polar, the charged strips have no effect on the stream of oil.

  The water molecules on the surface of water are so strong that the weight of a pin of needle cannot break them. For this reason, an object that is more dense than water can seem to float of the surface. When you exert an additional force with your finger and push on the needle, you break the surface tension and the needle sinks. Ethanol is also a polar molecule but has only one –OH and not two. Therefore, ethanol cannot create as many bonds as strong as those of water and thus has very little surface tension.

  The many strong hydrogen bonds of water also explain the observation that the water level can go above the level of the rim of a glass. Likewise, because ethanol has much lower surface tension, it cannot rise above the level of the rim.

  Liquid dish detergent is made of lipid molecules that have one end that is charged and another end that is very non-polar (uncharged). The charged ends interact with the water while the non-polar ends stand up away from the water thus forming a film that is one molecule thick. It spreads out very rapidly on the surface of water and pushes the pepper away.

Safety Precautions

- Make sure students are careful with sharp sewing needles and use them appropriately as directed in the procedure.

Answers to Analysis Questions

1. The stream of water is attracted to the acetate strip and the vinyl strip. The stream of oil was unaffected by either the acetate strip or vinyl strip. Therefore, the stream of water is attracted to both positive and negative charges. The oil is attracted to neither. Water must have negative charges that are attracted to the positively charged strip and positive charges that are attracted to the negatively charged strip. The oil has no apparent charges.
2. When the needles were touched, they sank to the bottom of the water. Therefore, the needles are denser than water. The needles floated on the surface of the water before the "skin-like" surface was broken. The positive and negative charges associated with the water molecules are strongly attracted to each other. A needle can appear to float because the water molecules are more strongly attracted to each other than they are to the atoms in the metal. You have to break these attractive forces before the needle will sink in the water.
3. The needles did not float on the ethanol. The attraction between ethanol molecules is not strong enough to support the weight of a needle.
4. The water rises above the rim of the glass in a dome-like shape because the positive and negative charges associated with water molecules are strongly attracted to each other and "hold each other together." The water spills only after the weight of the water above the rim of the glass is great enough to pull the water molecules apart.
5. Ethanol did not rise above the rim of the glass. It began to spill as soon as it reached the rim since the attractive forces between ethanol molecules are not strong enough to support any ethanol above the rim of the glass.
6. When the detergent was added to the pepper-covered water, the pepper instantly moved to the outer edge of the water surface. Since the detergent molecules have a charged head (in this case negative), they will be attracted to the water molecules. The long non-polar dipoles of the detergent molecules will protrude upward because they are attracted to each other but not to the water molecules. This breaks the surface tension in the centre. The remaining surface tension in the water pulls the pepper to the outside of the dish.

Answer to Conclusion Question

7. Water molecules are strongly attracted to each other resulting in a strong surface tension that can support small objects. Ethanol molecules are not as strongly attracted to each other and have a lower surface tension. Water molecules are polar since they were attracted to both positive and negative charges. From the structure of ethanol, C_2H_5OH, it can be deduced that ethanol molecules are also slightly polar, but much less so than water.

Assessment Options

- Use Assessment Checklists 2, 3, 4, and 5 as assessment aids.

Answers to Questions for Comprehension

Student Textbook page 69

Q5. Intramolecular forces are the forces within molecules that hold atoms together in molecules. Intermolecular forces are the forces between molecules that hold molecules together in the liquid or solid state.

Q6. Hydrogen bonds are similar to other types of dipole-dipole interactions because they are electrostatic forces between polar molecules, like all types of dipole-dipole interactions. Hydrogen bonds are distinct from other types of dipole-dipole interactions because they are stronger and they only occur when hydrogen is bonded to nitrogen, oxygen, or fluorine, the elements with the highest electronegativities. Hydrogen bonds occur only between hydrogen nuclei and atoms of these elements on an adjacent molecule.

Q7. London (dispersion) forces account for the state at any given temperature of many substances. They are the key force of attraction between non-polar molecules. As the numbers of electrons and the size of the positive nuclear charges increase, the state of the substances at room temperature goes from gas to liquid to solid. The family of halogen (Group 17) gases is an example of this phenomenon.

Section 2.2 Review Answers

Student Textbook page 70

1. Dipoles are molecules with oppositely charged ends separated by short distances and are found in polar molecular structures. Dipolar molecules tend to rotate so that the positive end of one molecule is nearest to and attracts the negative end of the next nearest molecule. Dipoles tend to be attracted together.
2. Hydrogen bonding applies to molecules where a hydrogen atom is covalently bonded to a highly electronegative atom. The only combinations where this occurs are H–O, H—N, and H—F.
3. Water has only hydrogen and oxygen atoms. The rigid structure of ice keeps the molecules further apart in ice than in liquid water, where the molecules move freely and the distances between molecules change constantly. However, each polar molecule in ice is surrounded by the oppositely charged ends of other polar molecules in a hexagonal pattern. This makes the bonds between molecules stronger and more rigid and opens up a regular pattern of spaces within the hexagon, which also accounts for the increased distance between the molecules.
4. A temporary dipole is a temporary positive or negative charge distribution within the molecule. This temporary charge on one end of a molecule causes a shift in charges in an adjacent molecule. This induces a dipole in the second molecule.
5. The two factors are the number of electrons in the molecules and the shape of the molecule. The more electrons and the larger the surface area, the stronger the London (dispersion) forces.
6. Intermolecular interactions determine the distance and the direction between molecules and the strength of the bonds between molecules in substances.
7. As intermolecular forces increase, molecules are held more tightly together and the state of the substance goes from gas to liquid to solid. The melting and boiling points are a measure of the strength of these forces. The stronger the intermolecular forces, the higher the melting and boiling points.
8. ICT Students may need help with the formatting and programming of a multimedia presentation. Become familiar ahead of time with the general attributes of the software they will be using so you can give required help. Assess the presentation for completeness and clarity of the concepts presented, and for originality and artistic.

2.3 Relating Structures and Properties

Student Textbook pages 71–80

Section Outcomes

Students will:

- relate structures of ionic lattices to their properties
- describe how an understanding of electronegativity contributes to the knowledge of melting and boiling points
- perform an investigation to illustrate the properties of ionic compounds
- analyze experimental data to determine the properties of ionic compounds
- relate melting and boiling points, and enthalpies of fusion and vaporization to predicted intermolecular bonding
- analyze data for trends and patterns on the melting and boiling points of a related series of molecular substances

Key Terms

melting point
heat of fusion
boiling point
enthalpy of vaporization

Chemistry Background

- Crystalline solids are classified as atomic solids (involving only London (dispersion) forces between atoms), molecular solids (involving all types of intermolecular forces between discrete molecules), network solids (involving covalent bonds holding particles together), and ionic crystals (involving electrostatic attractions between ions).
- The increase in London (dispersion) forces, which corresponds to increasing atomic mass, is really caused by

the increase in the number of electrons (and the corresponding increase in the number of protons) in more massive atoms.
- Melting point and boiling point can be used as a measure of strength of intermolecular forces, because if there were no intermolecular forces, all molecular substances would remain as gases due to the kinetic energy of the particles.
- All molecular substances have London (dispersion) forces. Dipole-dipole forces and hydrogen bonds are in addition to those forces.

Teaching Strategies

- Use concrete examples of each type of solid to illustrate their properties. Bring the samples to class and use them to illustrate the properties.
- Grow crystals during the discussion of ionic compounds to use in this section as examples.
- Do Investigation 2.E as a demonstration if there is not sufficient time to perform it as a laboratory activity.
- A number of overhead masters have been prepared for this section. You will find them with the Chapter 2 BLMs on the CD-ROM that accompanies this Teacher's Resource or at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs.
  Number (Type) Title
  2.3.1 (OH) Bonding in Solid, Liquid, and Gas Phases
  2.3.2 (OH) Melting and Boiling Points of Some Common Compounds
  2.3.3 (OH) Boiling Points of Binary Hydrides
  2.3.4 (OH) Conductivity of Ionic Compounds

Answers to Questions for Comprehension 8–10

Student Textbook page 75

Q8. The intermolecular forces in molecular compounds are generally much weaker than ionic bonds.

Q9. When molar mass increases, this means that there is an increase in the number of electrons and the size of the positive nuclei. As a result, the strength of intermolecular forces increase, holding the molecules of the solid more strongly together. Therefore the melting point increases.

Q10. Substances with hydrogen bonds have a higher melting point than substances stabilized by London (dispersion) forces. Since hydrogen bonds are the strongest intermolecular bonds and since London (dispersion) forces are part of every bond, hydrogen bonds can be thought of as an "add-on" force. Therefore, substances with hydrogen bonds would usually have a higher melting point than substances with only London (dispersion) forces.

Investigation 2.E: Properties of Substances

Student Textbook page 78

Purpose

Students will examine the properties of five different types of solids: non-polar covalent, polar covalent, ionic, network, and metallic. They will identify which substances fit into each of those five types by classifying them according to their properties (hardness, solubility, conductivity (in solution and in solid form), and melting point).

Outcomes

- 20–A2.1s
- 20–A2.2s
- 20–A2.3s
- 20–A2.4s

Advance Preparation

When to Begin	What to Do
3–4 weeks before	■ Ensure that all materials are available.
1 day before	■ Review Safety in your Chemistry Laboratory in the student textbook. ■ Review the procedure with students; have students begin the write-up. ■ Assign lab groups if you are using them. ■ Photocopy **BLM 2.3.5 (HAND) Investigation 2.E: Properties of Substances.**

Materials
■ 100 mL beaker and stirring rod ■ metal plate (iron or aluminium) ■ conductivity tester ■ ring stand with clamp ■ candle ■ Bunsen burner ■ match, lighter, or striker ■ timer ■ tongs ■ five unknown solids ■ distilled water

Time Required

45 minutes

Helpful Tips

- The following five substances work well for this investigation: paraffin wax (shavings and a block); sugar (granular sugar and a large solid piece such as rock candy); silicon dioxide (sand and quartz); salt (granular and crystal); tin (granular tin and tin foil).
- To conserve materials, set up three stations: one for testing hardness, one for testing electrical conductivity, and one for testing melting time. Alternatively, double the number of stations.
- To conserve resources, do the following: For solubility, tell students to use a single crystal, or only a few crystals, in a small test tube. For conductivity, provide solutions of the soluble substances (sugar and salt). For melting time, use only a few grains of solid each time.
- The metal plate can be made from a large can lid. Ask students to bring their own the day before, or start collecting lids ahead of time. You may need to burn off the thin layer of plastic on the inside and put that side down in the investigation.
- Use **BLM 2.3.5 (HAND) Investigation 2.E: Properties of Substances** to support this activity. Remove sections as appropriate to meet the needs of the students in your class. The answers to questions are on **BLM 2.3.5A (ANS) Investigation 2.E: Properties of Substances Answer Key.**
- ***Expected Results*** The exact results will vary with the specific samples chosen. In general, however, the metallic, ionic, and network solids will be harder and have the higher melting points. Polar covalent compounds will be more brittle and have lower boiling points and the non-polar covalent solids will be soft and have the lowest boiling points.

Answers to Analysis Questions

1. Sample table and answers for this choice of substances:

Solid #	**Hardness**	**Solubility in water**	**Solution conducts**	**Solid conducts**	**Melting point**
1 (paraffin)	softest	insoluble	—	no	melts first
2 (sugar)	second hardest	soluble	no	no	melts second
3 (quartz)	hardest	insoluble	—	no	does not
4 (salt)	fourth hardest	soluble	yes	no	does not
5 (tin)	third hardest	insoluble	—	yes	melts third

Paraffin melts easily, is soft, insoluble, and does not conduct electricity. It must have the weakest intermolecular forces, so it is non-polar covalent.
Sugar is quite soft, does not increase conductivity of water in solution, and is soluble. It has intermolecular forces only slightly stronger than paraffin. Therefore sugar is polar covalent.
Quartz is the hardest, but does not melt, is insoluble, and does not conduct electricity. The intermolecular forces are strongest of the substances. Therefore quartz is most likely a network solid.
Salt is also very hard, dissolves in water to form a conducting solution, but does not melt. The intermediate forces that hold salt together must be second strongest. Salt is therefore an ionic crystal.
Tin is third hardest, and conducts electricity. It also has a metallic lustre. It is most likely a metallic solid.

Answer to Conclusion Question

2. A non-polar covalent solid is a very soft, non-conducting solid that is insoluble in water and has a low melting point.
A polar covalent solid is a soft, non-conducting solid that is insoluble in water and has a low melting point.
An ionic solid is a hard, high-melting-point substance that is soluble in water. Although it will not conduct electricity as a solid, it does so in solution.
A metallic solid is a hard, high-melting-point substance that is insoluble in water and conducts electricity in the solid state.
A network solid is a very hard, high-melting-point substance that is insoluble in water and will not conduct electricity.

Connections (Nature of Science): Mane Products

Student Textbook page 79

Teaching Strategies

- Students will likely have favourite brands of hair-care products and be loyal to them for a variety of reasons. Have them research their favourites and be prepared to defend them using scientific criteria.
- Have students research hair-care products based on price, to make comparisons of very inexpensive, mid-price, and very expensive products.

Answers to Questions

1. A Google™ search using the search words "hair dyes" and "work" brings up many good sites. One of the first is **www.science.howstuffworks.com**. Other pages bring up sites for green and blue colours and other temporary colours, especially **www.spookyboutique.com**. Assess the answers by noting how temporary dyes work on a molecular level, and the different types (those that last for one shampoo, those that last for 12 to 24 shampoos).

2. Suggested criteria include: colour, attractiveness, positive associations with such things as beauty and popularity, and creativity. Have students ask several friends which packaging they liked and why.

3. A web search using the ingredients as search words will bring up sites that talk about safety as well as giving chemical information. However, refine the search by adding the search word "safety." Some of the information will be non-scientific. Much of it will not say specifically whether the ingredients are actually harmful in hair products. Students will need to draw conclusions from what they read. Assess answers by rating the completeness and clarity of the evidence given, whether the student draws a conclusion, and how well you think the conclusion follows from the evidence.

Section 2.3 Review Answers

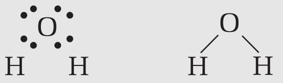
Student Textbook page 80

1. The state of a pure substance at room temperature is determined by the strength of the intermolecular forces holding it together.
2. **(a)** To melt an ionic compound, the ionic bonds holding the crystal lattice together must be broken.
   (b) To melt a metallic solid, metallic bonds must be broken.
   (c) To melt paraffin wax, a molecular solid, the intermolecular forces must be broken. Paraffin is a non-polar covalent compound, so London (dispersion) forces must be broken.
   (d) Since graphite is a network solid, covalent bonds must be broken when it is melted.
3. Students will likely say that chloroform, $CHCl_3^-(l)$, is polar since not all polar bonds are equal and they therefore do not cancel. In methane, CH_4, the polarities of all bonds are equal and thus they cancel. As a result, chloroform is slightly polar and the molecules will attract each other, giving it a higher boiling point than the non-polar methane. A more significant reason is that with the more massive chlorine atoms present in chloroform, there will be more attraction between the molecules regardless of polarity; for example CCl_4, though non-polar, is still a liquid at room temperature.
4. The higher the bond energy of a substance, the higher the temperature must be to overcome the forces of attraction between the particles in the substance. The ionic and metallic bonds must be broken in order for the compounds to melt or vaporize. The energy that breaks the bonds is the kinetic energy of particles. Temperature is a measure of the average kinetic energy of atoms, ions, or molecules. Therefore, high melting points and boiling points indicate that the bonding forces are strong. In molecular substances, covalent bonds do not break when a substance melts or vaporizes. It is the intermolecular attractions that must be broken to cause melting or vaporization.
5. The relatively low melting point of 140°C suggests that the solid is a molecular solid. The solubility of the solid indicates that it is polar. Therefore, the solid is likely a polar covalent solid. To confirm, test the solution of the solid for conductivity. Low or no significant conductivity would confirm a polar covalent solid. Test also for hardness, brittleness, malleability, and ductility.
6. Ionic compounds are hard and brittle. When an object made of ionic compounds is subjected to hammering or bending, the like charges of the particles can become aligned and repel each other. This will cause the crystal structure to break along the line where the charges are repelling. Therefore, ionic compounds are not practical compounds for tools.
7. **(a)** The melting point of the unknown solid is likely low. Since the compound does not conduct electricity as a liquid, it is a covalent compound.
   (b) As described in the answer to part (a), the compound contains covalent bonds.
8. Graphite is a solid of carbon atoms arranged as a network of hexagons layered in planes. The layers are held together by weaker bonds, so the electrons between the layers are held more loosely. When electricity is passed through the graphite parallel to the plane, these electrons in the area between the layers can move more easily, making the conductivity higher in that direction.
9. Assess the completeness and clarity of the laboratory design. Use Assesment Checklist 1.
10. Have students graph their data and mark the graph for correct labelling and following conventions. Melting and boiling points increase as the number of carbon atoms increases. This is because the boiling point and melting point of an alkane increases as the number of electrons increases. Larger molecules also have a greater surface area for Van der Waals attractions, so larger molecules have stronger attractions between them and therefore higher melting and boiling points. Smaller molecules have less surface area and fewer electrons, resulting in weaker Van der Waals interactions and lower boiling and melting points.

Chapter 2 Review Answers

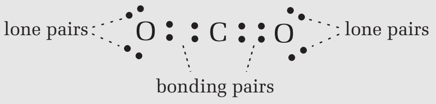
Student Textbook pages 82–83

1. **(a)**



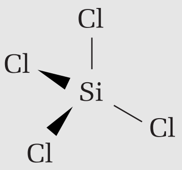
   Repulsion between the lone pairs (LPs) and the bonding pairs (BPs) causes the bent structure.

   (b)



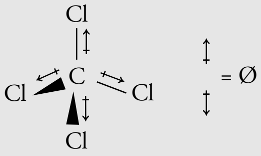
The oxygen LPs cause the oxygen atoms to be on opposite sides of the carbon atom (which has no LPs), resulting in a straight structure.

(c)



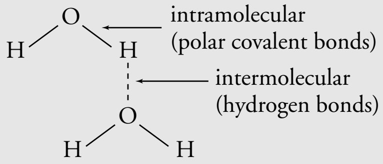
Mutual repulsion of BPs results in a tetrahedral arrangement.

(d)



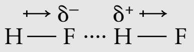
Although the bonds are polar, the polarities cancel out in this symmetrical molecule, so the molecule itself is non-polar.

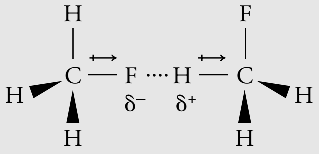
(e)



The molecules themselves are held together with polar covalent bonds (intramolecular bonds) between the hydrogen and oxygen atoms. The water molecules are held together with hydrogen bonds (intermolecular bonds).

(f)





In HF, since the hydrogen atom is bonded to the fluorine atom, hydrogen bonds will exist between molecules of hydrogen fluoride. In CH_3F, however, since the fluorine atom is bonded to the carbon and not the hydrogen atom, molecules of CH_3F will only be held together by dipole-dipole forces.

2. Oxygen, nitrogen, and chlorine are gases at room temperature because they are held together only by London (dispersion) forces. There is little attraction between the molecules because there are no dipoles; thus the molecules remain separate and the elements are gases. These elements also have relatively few electrons, compared to iodine, for example, which has more electrons and is solid at room temperature because the London (dispersion) forces are stronger.

3. Since ionic compounds do not contain molecules, it is not correct to talk about intermolecular forces when speaking of ionic substances. In ionic compounds, an ion is bound to several other ions of opposite charge in a lattice. Therefore in ionic compounds it would be correct to speak of forces between ions, or interionic forces.

4. All the halogen hydrides are polar, and atoms are held together by London (dispersion) and dipole-dipole forces. London (dispersion) forces are greatest in HI and smallest in HF. This is because the HI molecule is the largest and contains the most electrons. ΔEN is greatest in HF and smallest in HI, so the dipole-dipole forces are greatest in HF and smallest in HI. However, ΔEN is small for halogens, so the difference in the strength of the dipole-dipole forces is not as significant as the difference in the strength of the dispersion forces. Therefore, the boiling point should increase from HF to HI. In HF the relative sizes of the hydrogen atom and the fluorine atom allow hydrogen bonding to occur. Hydrogen bonding does not occur in the other hydrogen halides. As a result, HF has the strongest intermolecular force and its boiling point is higher than the boiling points of the other hydrogen halides.

5. Polar molecules are permanent dipoles. Non-polar molecules can have induced dipoles caused by nearby ions or polar molecules that distort the electron density of the non-polar molecules. The induced dipole will disappear when the charge entity moves away or the charge disappears. A polar molecule remains polar despite changes in the electric field around it.

6. The intermolecular forces of non-polar molecules are London (dispersion) forces. The strength of these forces depends on the number of electrons and the shape and size of the molecule. Oil is a mixture of large organic molecules that have long chains of atoms and a large number of electrons. As a result, the London (dispersion) forces between the molecules are very strong, even stronger than the hydrogen bonding in water. Therefore, cooking oil has a higher boiling point than water.

7. **(a)** London (dispersion) forces and dipole-dipole forces

(b) ionic bonds

(c) metallic bonds

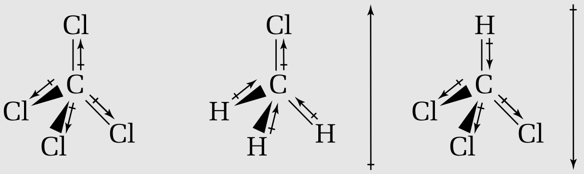
(d) London (dispersion) forces

8. Both $H_2O(\ell)$ and $NH_3(g)$ molecules have the same number of electrons. ΔEN is greater for O—H (1.24) than for N—H (0.84). Therefore, the O—H bond is more polar and forms stronger hydrogen bonds. Thus, hydrogen bonding is stronger in water than in ammonia.

This is also implied by the fact that water is a liquid at room temperature, whereas ammonia is a gas.

9. Order of increasing boiling points is $CH_3H_8 < C_2H_5OH < SiO_2$. SiO_2 is a network solid so it has the highest boiling point. CH_3H_8 and C_2H_5OH are covalent solids, which have lower boiling points than SiO_2. C_2H_5OH is polar, and it may have hydrogen bonding in solid form. C_3H_8 is non-polar and will not form hydrogen bonds. Therefore, C_2H_5OH probably has a higher boiling point than C_3H_8. Actual boiling-point data support this prediction: CH_3H_8 (231 K), C_2H_5OH (351 K), SiO_2 (2503 K).

10. The molecular shapes of CCl_4, CH_3Cl, and $CHCl_3$ are all tetrahedral.



CCl_4 is non-polar, but CH_3Cl and $CHCl_3$ would all have dipole-dipole forces between molecules in liquid samples. The order of the boiling points from lowest to highest is $CH_3Cl(\ell)$,) < $CHCl_3(\ell)$, and < $CCl_4(\ell)$. $CCl_4(\ell)$ has more electrons, so the London (dispersion) forces are stronger.

11. Dipole-dipole attractions are intermolecular forces between two polar molecules. An ionic bond is the attraction between two ions: one positively charged and the other negatively charged. Ionic bonds are usually thought of as intramolecular forces. In an ionic crystal, however, it is impossible to differentiate between intermolecular and intramolecular forces.

12. If the polar bonds in a molecule are arranged around the central atom in such a way that the dipoles can cancel each other, then the molecule is non-polar. Non-polar molecules with polar covalent bonds include binary (two-element) molecules with no LPs, such as CO_2, CCl_4, and BCl_3.

13. (a) PH_2F. Molecular shape is similar, so that is not a factor. Phosphorus has more electrons than nitrogen, so the London (dispersion) forces are stronger for the PH_2F. The electronegativity differences are greater between phosphorus and fluorine than they are between nitrogen and chlorine, so PH_2F is more polar and the dipole-dipole forces are greater.

(b) SiC. Silicon carbide forms a network solid, so the covalent bonds are very strong and would require very high temperatures to break. AsH_3 is a molecular substance held in the liquid state by weaker dipole-dipole and London (dispersion) forces.

(c) Xe. Xenon has more electrons than neon so its London (dispersion) forces are stronger.

(d) CH_3OH. CH_3OH has stronger hydrogen bonding than the other molecule.

(e) CH_3F. CH_3F has dipole-dipole forces but F_2 only has London (dispersion) forces.

(f) $AlCl_3$. The electronegativity differences are greater between aluminium and chlorine than they are between arsenic and chlorine, so $AlCl_3$ is more polar and the dipole-dipole forces are greater.
Note that $AlCl_3$ sublimes at 180 °C and does not boil. It also exists in the liquid form as Al_2Cl_6 dimers.

(g) C_4H_{10}. The chlorine molecule has the same number of electrons as the C_4H_{10} molecule and both are non-polar, but the C_4H_{10} molecule has a longer shape so its London (dispersion) forces are stronger due to the larger surface area.

(h) NH_3. NH_3 has hydrogen bonds, whereas PH_3 does not.

(i) C_4H_9F. C_4H_9F is more polar so it has stronger dipole-dipole forces.

(j) Ammonia. Ammonia has hydrogen bonding.

(k) Silicon dioxide. Silicon dioxide is a network solid.

(l) Krypton has more electrons and so its London (dispersion) forces are stronger.

(m) KCl. KCl is ionic.

(n) If the formula contains a metal and a non-metal such as Mg and O, the compound is ionic. If the formula contains only non-metals such as N and H, the compound is molecular.

(o) PH_2Cl. PH_2Cl has more electrons, so the London (dispersion) forces are greater.

(p) NH_4Cl. NH_4Cl is an ionic solid.

(q) C_2H_5F. C_2H_5F is more highly polar.

(r) CH_3NH_2. CH_3NH_2 is a longer molecule, so the London (dispersion) forces are greater.

(s) C_2HCl. C_2HCl is polar, so there are dipole-dipole forces.

(t) H_2O. H_2O has strong hydrogen bonding.

(u) SiO_2. SiO_2 is a covalent solid.

(v) NaCl. NaCl is an ionic solid.

14. (a) Glycerol molecules are larger and longer than water molecules, so they would have stronger London (dispersion) forces. They also have three O—H bonds per molecule so there is a greater opportunity for hydrogen bonding. Glycerol has stronger bonds and therefore flows much more slowly.

(b) $CaCl_2$ is an ionic compound and is not an individual molecular unit.

(c) Silver forms metallic bonds and is malleable because the sea of electrons allows planes of metallic ions to slide past each other.

(d) Graphite is a network solid and so is fairly stable at high temperatures. The layers of carbon slide over each other easily, so it is an effective lubricant.

(e) In diamond, all the electrons are localized in covalent bonds and cannot move.

(f) In the molten substance, the rigid lattice structure is disrupted and the ions are free to move.

(g) The varying charges on ions result in variable melting points.

(h) The regular arrangement of carbons in a crystalline structure and the strength of the covalent bonds imply that diamond is the hardest substance.

(i) Ice has a crystalline structure that keeps the water molecules farther apart in ice than in liquid. As the water freezes, it expands, breaking the bottle.

(j) Graphite is a strong but light-weight network solid.

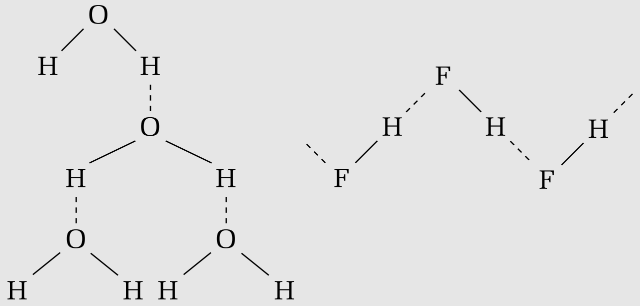
Answers to Applying Concepts Questions

15. Measure the melting point of each substance. The ionic compound should have the higher melting point. Then test the electrical conductivity of the melted substances. The ionic material should conduct electricity, while the covalent material should not. Test the solubility of the materials in water. If they dissolve in water, test the conductivity of the solution. The ionic material will conduct electricity in solution, while the covalent material will not.

16. Liquid A is polar since it will move towards a charged rod, indicating an opposite charge on one end of the molecule. Liquid B is non-polar since it is unaffected by a charged rod.

17. As water cools toward 4 °C, the molecules move more slowly and the attractive forces pull them closer together, so the density of the water increases. This increase in density continues until the temperature of the water is 4 °C. At 4 °C the open hexagonal structure (due to hydrogen bonding) begins to form, stopping further contraction in the volume of the water. When the water cools further, below 4 °C, the water molecules form more of the open hexagons that are found in ice. The volume begins to increase, so the density begins to decrease.

18. From electronegativity values, the H—F bond ($\Delta EN = 1.78$) is more polar than the O—H bond ($\Delta EN = 1.24$). This suggests that hydrogen fluoride may have a higher boiling point than water. However, water has two polar bonds whereas hydrogen fluoride has only one. The bent shape of the water molecules allows for two O—H bonds to form hydrogen bonds. Each water molecule can be linked to as many as five other molecules, but each hydrogen fluoride molecule can be linked to only two other molecules. Water boils at 100 °C and hydrogen fluoride at 19 °C. (Students would be advised to use search words "hydrogen fluoride boiling point"; using HF gives sites for the element hafnium as well.)



19. SF_2 is a bent molecule and therefore is polar because the S—F bonds are polar. SiF_4 is tetrahedral molecule. The dipoles of the tetrahedron cancel each other, making the molecule non-polar.

20. The electronegativity difference between Na and I is 1.73. NaI is an ionic solid. Therefore, the forces of attraction that must be overcome to melt this solid are electrostatic attractions due to ionic bonding. The electronegativity difference between H and I is 0.46. HI is a covalent solid that is only slightly polar. The intermolecular forces in HI are London (dispersion) forces and dipole-dipole forces which must be overcome to melt HI. These forces are much less than those for NaI.

21. **(a)** Diamond is a crystalline solid that does not melt. It sublimes at a temperature around 4000 K. It is one of the hardest substances known. 2,2-dimethylpropane is a liquid at room temperature. Its melting point is 257 K, and its boiling point is 283 K.

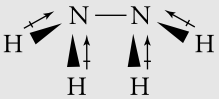
(b) The C—C bonds in both diamond and 2,2-dimethylpropane are the same. In both substances, a carbon atom is bonded to four other carbon atoms in the shape of a tetrahedron. In diamond, however, each carbon atom is bonded to four other carbon atoms in a huge network. In 2,2-dimethylpropane, each outer carbon is bonded to one carbon atom and three hydrogen atoms. Thus, 2,2-dimethylpropane exists as discrete molecules. The diamond molecule is the diamond itself. The entire crystal is a molecule because every carbon atom in the crystal is linked to four other carbon atoms within the network.

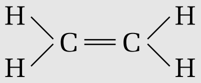
(c) The C—C bonds in diamond are networked. To cleave, to smash, or to scratch a diamond, you have to break thousands of C—C bonds. This makes diamond a very hard solid. For 2,2-dimethylpropane, the forces that hold the molecules in a solid are London (dispersion) forces. Thus, the solid is much softer and melts at a much lower temperature.

22. Carbon dioxide is a linear, non-polar covalent molecule. The only intermolecular forces that act between the molecules are London (dispersion) forces. These forces are very weak because the carbon dioxide molecule has only 22 electrons and is quite small. Therefore, carbon dioxide exists as free molecules (a gas) at room temperature. In contrast, each silicon atom is strongly bonded to four oxygen atoms in the form of a tetrahedron, and each oxygen atom is bonded to two silicon atoms to form a network. To melt silica, a lot of energy is needed to break the bonds in the network. Therefore, silica has a very high melting point (1873 K), and it exists as a solid at room temperature.

23. The ΔEN of the Se=O bond is 0.89. Thus each Se=O bond is polar. The shape of the SeO_2 molecule is bent, and a net dipole results. Therefore, the molecule is polar. In SeO_3, the selenium atom forms three double bonds with three oxygen atoms. The shape of the molecule is trigonal planar. Although the Se=O bonds are polar, the symmetry of the molecule causes the dipoles of the Se=O bonds to cancel one another. Therefore, the SeO_3 molecule is non-polar.

24. Ethene, C_2H_4, is a planar molecule with a double bond between the two carbon atoms. In this arrangement, the slight polarities of the C—H bonds cancel each other. Therefore, ethene is a non-polar molecule. The intermolecular forces in ethene are London (dispersion) forces only. Hydrazine, N_2H_4 (18 electrons), has stronger dispersion forces than ethene (14 electrons). The N—H bond is also more polar than the C—H bond. A hydrazine molecule consists of two trigonal pyramidal structures linked through the nitrogen atoms. In this arrangement, the polarities of the N—H bonds do not cancel, and the molecule is polar. As well as dipole-dipole forces and London (dispersion) forces, hydrogen bonding occurs between the hydrogen atoms of one molecule and the nitrogen atoms of another. Therefore ethene would have a lower boiling point.





25. Argon is a noble gas, and fluorine gas is a non-polar covalent molecule. Both of these substances would be held together by London (dispersion) forces. Since they have the same number of electrons, the strength of the London (dispersion) forces would be similar, so they should have similar boiling points.

26. Krypton gas atoms would be held together with London (dispersion) forces. Since London (dispersion) forces are the weakest, the boiling point of krypton should be the lowest. Hydrogen bromide gas is a slightly polar molecule, so the gas would be held together with weak dipole-dipole forces. Hydrogen selenide gas is also slightly polar and would be held together with weak dipole-dipole forces. The boiling points of HBr(g) and $H_2Se(g)$ should be similar, since they are held together with similar forces.

27. **(a)** network
(b) ionic
(c) molecular
(d) metallic

Answers to Making Connections Questions

28. There are no correct or typical answers to questions such as this. However, students should include the following ideas and provide reasons for their opinions. *Risks associated with inert substances:* Materials that are very inert present problems with disposal. A common example is foam cups, which do not decompose and therefore end up in landfill sites. If every 12-year-old bought a Kevlar^{TM} jacket for protection, imagine the number of outgrown Kevlar^{TM} jackets in landfill sites three years later.
Benefits associated with inert substances: Kevlar^{TM} gloves and vests protect workers in hazardous situations. Foam cups are inexpensive and convenient for keeping drinks warm or cold.
Benefits versus risks: The benefits of Kevlar^{TM} products outweigh the risks, since the quantity in use remains quite small. The benefits of foam cups do not outweigh the risks, since people can use re-usable or biodegradable cups instead.

29. Some examples of how intermolecular forces influence the weather: If you are standing at the top of a ski hill and snow just begins to fall, you may feel the temperature rising. As water sublimes to form snowflakes, hydrogen bonds (intermolecular forces) are formed. The making of bonds produces energy, so you may feel the air temperature rising. You get out of a swimming pool on a very hot summer day and lie on a hot concrete patio beside the pool. You find yourself shivering. As the water on your body evaporates, the hydrogen bonds in the water need energy in order to break, and absorb heat from your body, making you cooler.

On a cold winter morning, you may notice frost on classroom windows. During the previous night, the glass window became very cold. As water vapour passed over the window, it cooled and the water molecules were brought closer together. Hydrogen bonds formed to make small crystals of ice. The frost "painting" on the window was formed by water vapour subliming to the solid. Because of the tremendous amount of heat necessary to change the state of water, it has a moderating effect on the temperatures. The average of all winter and summer temperatures for Edmonton is 2 °C, for Fort Chipewyan it is –2 °C, for Calgary it is 3 °C, and for Medicine Hat it is 5 °C. Vancouver has the highest average in Canada, around 10 °C.

30. **(a)** Although atoms do not become charged as do clouds, a separation of charge may form in an atom due to electron movement. As with a cloud, the negative "part" of an atom will lead to a separation of charge in nearby atoms, just like the negatively charged cloud causes a positive leader. The opposite charges in the atoms lead to the attractive London (dispersion) force. Similarly, the opposite charges in the cloud and in the positive leader result in the leader moving (being attracted) to the cloud.

(b) The separation of charge in an atom leads to a separation of charge in nearby atoms and the resulting London (dispersion) force. This is similar to what happens when people are near a charged cloud. The charged cloud leads to a separation of charge in the nearby peoples' hair, causing an attraction between the hair and the cloud, and causing their hair to stand on end.

Career Focus: Ask a Nanotechnologist

Student Textbook 84–85

Teaching Strategies

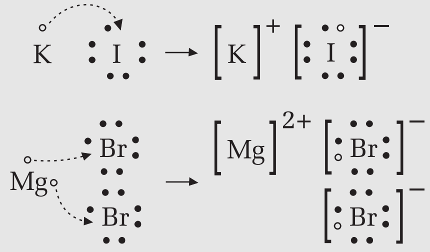
- Nanotechnology has been heralded as "the next industrial revolution," with the potential to cause as much disruption—both economic and social—as the first one did, while promising to deliver almost unimaginable benefits.
- Canada's National Institute for Nanotechnology is an integrated, multi-disciplinary institution involving researchers in physics, chemistry, engineering, biology, informatics, pharmacy and medicine can be accessed at http://nint-innt.nrc-cnrc.gc.ca/home/index_e.html
- There are growing calls for consideration of all aspects of nanotechnology. The Center for Responsible Nanotechnology is a US-based organization with global links that looks at both good and bad social implications for nanotechnology and highlights the need for public policy to govern the field. Nanotechnology news, discussions of related issues, and other links can be found at http://www.crnano.org/news.htm. The site's section on "Nanotechnology Basics For Students and Other Learners" raises some disturbing issues about the potential use for weapons of mass destruction. Students venturing beyond the research suggested in the feature may find information related to this issue on their own and may wish to raise the topic in class.

Answers to Go Further… Questions

1. Nanoscience takes matter and creates matter structures that are different from matter in traditional solid, liquid, or gas states. These new structures also have properties that are different from the traditional structures.
2. Silicate emits light poorly, but silicate nanoparticles are great light emitters, illustrating the concept that substances exhibit different properties on a nanoscale than they do under normal conditions.
3. Several good sites exist for searches. One is **http://www.azom.com/news.asp?newsID=1873**, which links to **http://www.azom.com/searchresults.asp?matkeyword=nanomaterial.**
4. Good search words include *nanotechnology* and *drug delivery*, and *nanomedicine*. Nanotechnology systems can deliver drugs to tiny areas within the body, eliminate side effects of traditional drug therapies, permeate cell walls, and be more efficient and effective. The examples chosen should illustrate these advantages.

Unit 1 Review Answers

Student Textbook 86–89

1. **(b)**
2. **(e)**
3. **(c)** and **(d)**
4. **(c)**
5. **(e)**
6. **(a)** and **(c)**
7. **(c)**
8. **(d)**
9. **(b)**
10. **(b)**
11. The three properties of ionic and molecular compounds that differ are boiling temperature, electrical conductivity, and hardness. For ionic compounds: high boiling points, high electrical conductivity in liquid state, hard, and brittle. For molecular compounds: low boiling points, low electrical conductivity, soft.
12. In electron dot diagrams of atoms, some electrons are drawn in pairs when the atom has more than four valence electrons. This corresponds to the actual placement of electrons in energy sublevels in atoms. Each sublevel can hold two electrons, and until all four sublevels are half-filled, one electron is placed into each of the sublevels. When more than four valence electrons are present, pairs will form in the energy sublevels.
13. Most ionic compounds are composed of both metals and non-metals. Molecular compounds usually contain only non-metals.
14. 

15. **(a)** nickel(II) oxide
    (b) sulfur trioxide
    (c) potassium carbonate

(d) tetraphosphorus decaoxide

(e) sodium perphosphate

(f) ammonium hydrogen carbonate

(g) dinitrogen monoxide

(h) sodium hypochlorite

(i) silicon dioxide

16. (a) HgS(s)

(b) $Ca(OH)_2(s)$

(c) $Cl_2(g)$

(d) SO2(g)

(e) CsF (s)

(f) $NH_4NO_2(s)$

(g) MnO2(s)

(h) $Pb(SO_3)_2(s)$

(i) CaOOCCOO(s)

17. Intramolecular forces exist between atoms in a molecule, such as the covalent bonds between the hydrogen and oxygen atoms in water. Intermolecular forces exist between molecules, such as the hydrogen bonds that hold water molecules together.

18. Predicted bond angles are based on maximum separation. In a tetrahedral, the maximum angle that can be formed between bonds is 109.5°. In a trigonal planar molecule, that maximum is 120°. These theoretical bond angles would be seen if all of the atoms in a molecule (and therefore the repulsions) were equal. But, lone pairs repel each other to a greater degree than a lone pair-bonding pair and even more so than bonding pairs, so the actual bond angles are affected by the differing degrees of repulsion.

19. (a) Alkanes are tetrahedrals.

(b) They are symmetrical molecules, so they are non-polar. London (dispersion) forces would exist between the molecules.

(c) Since alkanes are symmetrical, they are non-polar.

(d) Since they all contain the same number of electrons, one could expect that their boiling points would be similar. In reality, however, the boiling point depends not only on the number of electrons, but also on the efficiency of molecular packing. The better they pack, the higher the boiling point.

20. (a) Radon and krypton are both noble gases, held together only by London (dispersion) forces. Since radon has more electrons, it will have a higher boiling point.

(b) Hydrogen chloride, HCl(g) is a polar molecule, so molecules of HCl(g) are held together with dipole-dipole forces, which are stronger than the London (dispersion) forces holding atoms of argon together.

(c) Methane, $CH_4(g)$, is a tetrahedral molecule. Although the individual bonds are polar, the dipoles cancel in this symmetrical molecule, making methane non-polar.

(d) To determine the shape of hydrogen sulfide, consider the bonding elements around sulfur. Sulfur has two lone pairs and will form two single bonds with the hydrogen atoms. This is VSEPR class AX_2E_2, a bent molecule.

21. Lone pairs of electrons around the central atom of a molecule will reduce the bond angle between two bonding pairs of electrons.

22. In ionic compounds there are no distinct pairs or groups of atoms that can be identified as “molecules.” Due to their opposite charge, an attractive force exists between every negative ion and every positive ion that surrounds it, and vice versa.

23. Chlorine (Cl_2) consists of two chlorine atoms that share an electron pair. Each individual chlorine atom has seven valence electrons before bonding. Oxygen (O_2) molecules consist of two oxygen atoms that share two electron pairs. Each individual oxygen atom has six valence electrons. Nitrogen (N_2) molecules consist of two nitrogen atoms that share three electron pairs. Each individual nitrogen atom has five valence electrons. The double and triple bonds form to complete a stable octet for each atom.

24. A special model was needed to explain the unique properties of metallic compounds, which are very different from those of molecular compounds.

25. CO_2 and CS_2 are linear non-polar molecules. Dispersion forces are the only intermolecular forces that determine their melting points and boiling points. Molecules of CS_2 (38 electrons) are larger than molecules of CO_2 (22 electrons). Therefore, London (dispersion) forces are stronger in CS_2, and its boiling point is higher than the boiling point of CO_2.

26. The difference in electronegativities can be used to predict the type of bond that will form between two atoms of any two elements. If the difference is greater than 1.7, the bond will most likely be ionic. If it is less than 0.5, the bond will likely be covalent and slightly polar. If the difference falls between 0.5 and 1.7, the bond is polar covalent. If there is no difference, the bond will be non-polar covalent.

27. The bond would be expected to be strong and cause the solid to form an ionic crystal lattice with a high melting point. The bond would be 30% covalent.

28. Dipole-dipole attractions are between molecules that have a permanent positive charge on one end and a negative charge on the other end. London (dispersion) forces are between molecules that have a temporary dipole caused by vibration of electrons, setting up an uneven distribution of charge. For a small molecule, such as hydrogen chloride, the dipole-dipole forces are stronger than the dispersion forces. However, dispersion forces can be quite large if there are a large number of electrons spread over a large molecule.

29. In graphite, each carbon atom forms three strong covalent bonds with three other carbons in a trigonal planar pattern. These bonds give a layered structure. The fourth bond, between the layers, is longer and weaker than the other three. It is easily broken and allows the layers to slide by one another, making it a good lubricant. Because of the strong covalent network bonds, graphite does not break down at high temperatures as most lubricants do.

30. *Similarities:* Both generally have high melting temperatures. Both are hard but covalent network solids are harder. Both involve bonding between atoms. Both generally have a high boiling point and are insoluble in water.
Differences: Metals are malleable and ductile and network solids are brittle. Metals conduct electricity and network solids are insulators.

31. When pure molecular substances boil, the intramolecular covalent bonds between the atoms do not break. It is the intermolecular bonds between the particles that must break, and, therefore, the boiling point is an indication of the strength of intermolecular attractive forces.

32. The strength of dispersion factors is affected by the number of electrons in the molecule and the shape of the molecule.

33. Hydrogen bonds require a hydrogen atom bonded to a highly electronegative atom, such as oxygen, nitrogen, or fluorine, and to a lone pair of a nearby electronegative atom.

34. Hydrogen bonds have their maximum strength when they are oriented along a line from the centre of one oxygen atom, through the covalently bonded hydrogen atom, and to the centre of the oxygen atom of the next water molecule. This occurs in ice.

35. Water forms hydrogen bonds between the hydrogen atoms in one water molecule and the oxygen atom in another. Methane does not form hydrogen bonds, and, therefore, has a lower boiling point. Ammonia forms hydrogen bonds between the hydrogen atoms in one ammonia molecule and the nitrogen atoms in another. However, the bonds between nitrogen and hydrogen atoms in ammonia are less polar than O—H bonds, and, therefore, ammonia has a lower boiling point.

36. Metals are malleable because the layers in their structure can move without breaking the array, since freely moving valence electrons continue to exert a uniform attraction on the positive ions in the layers.

37. **(a)** The shape of ionic crystals is determined by the relative size and charge of the ions in the crystal.

(b) Ionic crystals do not conduct electricity (unless dissolved or melted) and are brittle. They have smooth sides and a characteristic shape. When struck, the crystal may cleave cleanly into smaller crystals of the same shape.

(c) In order to conduct electricity, there must be freely moving charged particles. In a crystal, the ions are fixed, so cannot conduct electricity. The remainder of the properties are due to the structure of the crystal. Ions will pack into a crystal shape based on the relative size and charge of the ions, leading to a specific shape. It must be neutral in charge, so the faces will be smooth. When struck, the like-charged ions will be forced beside each other, leading to the cleaving of the crystal.

38. In order to melt a substance, the intermolecular forces holding the molecules/atoms or ions together must be broken. The stronger the intermolecular forces, the greater the amount of energy that must be put in to melt the substance. This amount of energy is measured as the kinetic energy in the molecules, or the melting point.

39. In the molten state, ions are free to move and can act as charge carriers. However, in the solid state, ions are held in rigid positions in the crystal lattice and cannot move to carry an electric charge.

Answers to Applying Concepts Questions

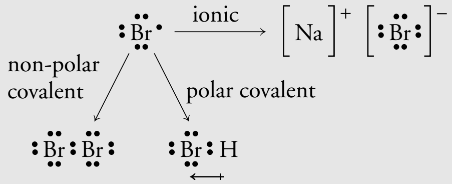
40. **(a)** mostly ionic

(b) polar covalent

(c) metallic

(d) non-polar covalent

41.



42. The electronegativity difference for LiBr is 1.98 and for LiI is 1.68. LiBr should have more ionic character than LiI. Thus, LiBr should have a higher melting point than LiI.

43. Electronegativities: N (3.04), H (2.20), F (3.98), I (2.66), Cl (3.16)

(a) N is more electronegative, and it is the negative end of the dipole (N ← H).

(b) F is more electronegative, and it is the negative end of the dipole (F ← N).

(c) Cl is more electronegative, and it is the negative end of the dipole (I → Cl).

44. Electronegativities: F (3.98), Cl (3.16), Br (2.96), I (2.66), Si (1.50), P (2.19), S (2.58)

(a) The order of increasing bond polarity is Cl—Cl, Br → Cl, Cl → F. The arrow points toward the element with a partial negative charge. Cl—Cl is non-polar.

(b) The order of increasing bond polarity is Si—Si, S → Cl, P → Cl, Si → Cl. The arrow points toward the element with a partial negative charge. Si—Si is non-polar.

45. First determine whether the solid can conduct electricity. If it can conduct electricity as a solid, it is a metallic compound. You can confirm this if it has lustre and is malleable and ductile. If the solid can conduct an electric current when it is dissolved in water, it is an ionic compound. If the compound cannot conduct electricity, heat the solid to determine the melting point. A high melting point will indicate a network solid. If the boiling point is low, the compound is molecular.

46. The ammonium ion is symmetrical, having a tetrahedral shape, and is therefore non-polar because the bond polarities cancel each other. Since the ammonia molecule has only three hydrogen atoms, it forms a pyramidal shape and is not symmetrical. Therefore, the bond polarities do not cancel each other and the molecule is polar.

47. Agree. All molecules have random motion due to their kinetic energy. If there were no intermolecular forces to hold them together, collisions with each other would bounce them apart and they would remain as gases. It is only the intermolecular forces that hold them together in the liquid or solid state.

48. Boiling point generally increases down a group, because as the number of electrons increases, the strength of the London (dispersion) forces increases. The expected trend in boiling points is $H_2O < H_2S < H_2Se$. However, H_2O does not follow the trend because there is hydrogen bonding in H_2O, but not in the others.

49. Covalent bonds are the simultaneous attractions between the nuclei and electrons of two or more atoms.

(a) In non-polar covalent bonds, the attractions of the different nuclei of the atoms for the electrons are the same.

(b) In polar covalent bonds, the nucleus of one of the atoms attracts the bonding electrons more strongly than the nuclei of the other atom does. In the O—H bond in a water molecule, the oxygen nucleus attracts the pair of electrons between the two atoms more than the hydrogen nucleus does. Therefore, the O—H bond in water is polar covalent.

(c) In ionic bonds, ions of opposite charges attract each other. Particles in ionic compounds cannot properly be termed molecules, so water cannot be ionic.

50. **(a)** London (dispersion) forces are temporary attractive forces that form between all molecules due to an imbalance in charge that induces an opposite charge in nearby molecules. They are the only intermolecular force that exists between non-polar molecules and noble gases.

(b) Dipole-dipole forces are attractive forces that exist between the oppositely charged ends of polar molecules.

(c) Hydrogen bonds are extremely strong dipole-dipole forces that exist between molecules where hydrogen is bonded directly to either oxygen or fluorine or nitrogen. This best describes the forces between water molecules in a glass of water.

(d) Intermolecular forces are attractive forces that exist between atoms, molecules, or ions. They must be broken in order to change the state of the substance.

(e) Intramolecular forces exist between atoms within the same molecule.

51. **(a)** A metallic element has metallic lustre, is a good conductor of electricity and heat, is malleable and ductile, and melts at a relatively high temperature.

(b) A metallic element reacts with a non-metallic element to form an ionic compound. It reacts with an acid to produce hydrogen gas. (Of course only active metals will do this, but this is a reaction with which students should be familiar. A metallic element also reacts with oxygen to form oxides that dissolve in water to form strong bases. It reacts with chlorine to form salts that conduct electricity in the molten state or in an aqueous solution.)

52. **(a)** The percent ionic character increases with increasing ΔEN. Thus, the greater the electronegativity difference between the elements, the more ionic character a bond has.

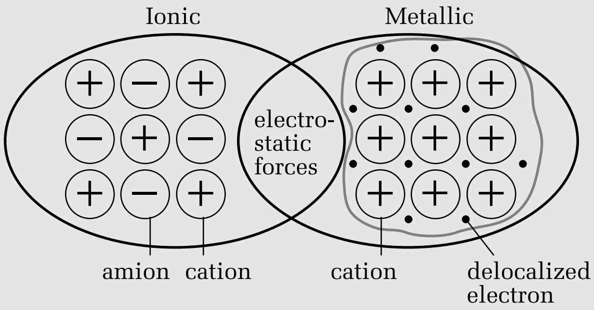
(b) Chlorine, Cl_2 , has 0% ionic character. This implies that the two chlorine nuclei attract the pair of electrons between them equally. The bond is described as a non-polar covalent bond.

(c) The trend of the graph suggests that there is no compound with 100% ionic character. This implies that the model for covalent bonding applies to all compounds. Even for ionic compounds with large electronegativity differences, complete electron transfer (as suggested by the ionic bonding model) is unlikely.

(d) The arbitrary cutoff suggests that bonding is a continuum with the ionic bond on one extreme and the non-polar covalent bond on the other extreme. Most of the bonds are somewhere in between and have at least a partial polar covalent nature. (Thus, all bonds are considered to be the simultaneous attractions between the nuclei and electrons of two or more atoms.)

53. Fluorine has the highest electronegativity and so forms very strong bonds with other compounds. Therefore, from a bonding point of view, fluorine is difficult to isolate. It is also extremely poisonous. (Once prepared, the fluorine gas attacks water and vessels. It cannot be contained, and it reacts with moisture in the air to form deadly hydrogen fluoride gas. The poisonous nature of hydrogen fluoride, which injured many people who worked with the gas in the past, makes the safe isolation of fluorine more remarkable. Moisson himself reported four delays in his studies of the electrolysis due to severe poisoning. He captured the fluorine gas in a platinum-iridium alloy U-tube, which was cooled to reduce the rate of attack by fluorine on the vessel walls. Thus, the Nobel Prize was given to Moisson for his dedication and courage working with extremely toxic materials to forward the cause of science.)

54.



Answers to Solving Problems Questions

55. **(a)**

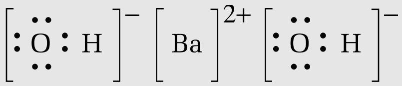
(b)

(c)

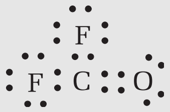
(d)

(e)

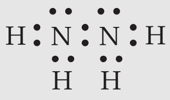
56.



57. **(a)**



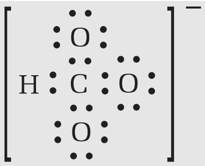
(b)



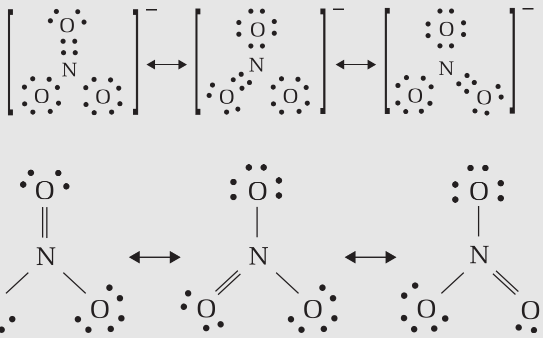
(c)

(d)

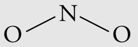
(e)



(f)

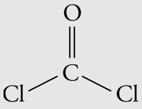


58. **(a)**



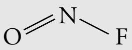
(bent)

(b)



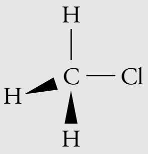
(trigonal planar)

(c)



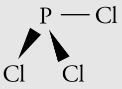
(bent)

(d)



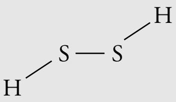
(tetrahedral)

(e)



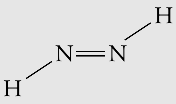
(pyramidal)

(f)



(bent)

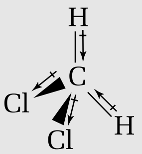
(g)



(bent)

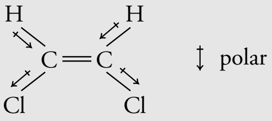
59. a, b, c, d and e are all polar.

60.

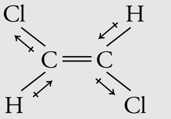


The two atoms of chlorine and the two atoms of hydrogen are arranged in a tetrahedral shape around the carbon. The C—Cl bonds are more polar than the C—H bonds. The direction of the polarities indicates that the Cl side of the molecule is slightly negative and the H side is slightly positive, so the molecule is polar.

61.



polar



non-polar

This molecule is trigonal planar. There are two isomers; one is polar, the other is non-polar.

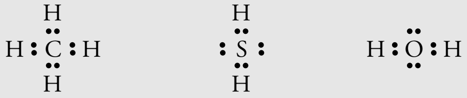
Answers to Making Connections Questions

62. Student answers will depend on the type of research and application selected. Possible applications include drug delivery, protective films, superconductors, and nanotubes. Students may be divided into groups to research specific applications. Each group can then present their information to the class.

63. **(a)** Iron is held together with metallic bonds. The properties of metallic bonds include ductility and malleability. This allows iron to be used in many applications that require strength and some flexibility. Iron(III) oxide is an ionic compound. Although it will have a high melting and boiling point, ionic compounds are brittle and can be easily cleaved into smaller pieces.

(b) One possible difference is that the ratio and size of the ions is different between iron(III) oxide and aluminium oxide. This would lead to a different crystal structure and perhaps different properties.

64. **(a)**



methane

hydrogen sulfide

water

(b) The intermolecular forces holding these molecules together differ, and so will their boiling points. Methane is non-polar, held together only by the weakest intermolecular forces, London (dispersion) forces, so methane will have the lowest boiling point. Water is held together by the strongest intermolecular forces, hydrogen bonds, so water will have the highest boiling point. Hydrogen sulfide gas is polar and held together with dipole-dipole forces, so it will have a boiling point between the other two substances.

(c) There are two factors affecting the boiling points of these substances. The polarity of the molecules is slightly less, so the dipole-dipole forces will be reduced somewhat; however there are more electrons in H_2Te than in H_2Se (and even H_2S), which will lead to stronger London (dispersion) forces. In reality, the boiling point of H_2Te is greater than H_2Se, which is greater than H_2S.

(d) Using the differences in boiling points, the sour gas could be cooled. When the temperature falls below the boiling point of hydrogen sulfide, it can be removed as a liquid, leaving only pure methane behind.

65. **(a)** To separate the strands of DNA, the DNA could be heated until the hydrogen bonds are broken. This would not break the covalent bonds in the strands.

(b) If the temperature of the protein gets too high, the hydrogen bonds in the protein will be broken, changing its shape. Since the functionality of a protein is based on its shape, if the temperature gets too high, the protein will no longer function.