Unit 4 Atomic Structure, Periodic Trends, Molecular Geometry, Bonding & Intermolecular Force - Stretching

Content Area:	**Science**
Course(s):	**AP Chemistry**
Time Period:	**December**
Length:	**6 weeks**
Status:	**Published**

Transfer Skills

Review of basic concepts from first year chemistry, atomic structure, periodic table, molecular geometry, bonding, & intermolecular forces, with additional focus on the depth required for Advanced Placement curriculum.

Enduring Understandings

Each electron has its own distinct amount of energy.

Understanding of regularities and patterns in the periodic table allows for predictions of interactions among the elements.

An atom’s electron configuration determines who it interacts with to form bonds.

The Lewis dot structure and electron domain placement determine the shapes and therefore the properties of molecules.

Strength of the intermolecular forces affects physical and chemical properties.

Essential Questions

How does the structure and composition of the atom influence its chemical and physical properties?

To what extent are all atoms different and / or same?

How does the electron behave?

How can properties of elements be predicted using the periodic table?

How does chemical bonding determine the properties of a substance?

How does the molecular structure influence the properties of a substance?

Content

Bond order, formal charges, paramagnetic, diamagnetic, allotropes, bond enthalpy, cohesive, adhesive, ionization energy, isotope, electron affinity, resonance, VSEPR, quantum numbers, electron configuration, emission spectra

Skills

Deduce and infer atom structure data from the periodic table

Calculate the energy difference between any two allowable energy states of the electron.

Write the electron configuration of elements.
Explain how atomic radii, ionization energy, and electron affinities vary within a group and within a period on the Periodic Table

Differentiate between the properties of polar covalent, non-polar covalent, & ionic bonds including intermolecular forces ion-ion force, dipole, hydrogen bonding & dispersion forces

Draw electron dot structures for simple covalent molecules containing single, double and triple bonds and resonance.

Explain why a liquid has a vapor pressure and why a change in temperature causes a change in vapor pressure.

Interpret vapor pressure diagrams, phase diagrams, and heating cures.

Resources

Standards

Big Idea 1: The chemical elements are fundamental building materials of matter, and all matter can be understood in terms of arrangements of atoms. These atoms retain their identity in chemical reactions. Enduring understanding 1.A: All matter is made of atoms. There are a limited number of types of atoms; these are the elements. Essential knowledge 1.A.1: Molecules are composed of specific combinations of atoms; different molecules are composed of combinations of different elements and of combinations of the same elements in differing amounts and proportions.
a. The average mass of any large number of atoms of a given element is always the same for a given element.

Enduring understanding 1.B: The atoms of each element have unique structures arising from interactions between electrons and nuclei. Essential knowledge 1.B.1: The atom is composed of negatively charged electrons, which can leave the atom, and a positively charged nucleus that is made of protons and neutrons. The attraction of the electrons to the nucleus is the basis of the structure of the atom. Coulomb's Law is

qualitatively useful for understanding the structure of the atom.
a. Based on Coulomb's Law, the force between two charged particles is proportional to the magnitude of each of the two charges ($q1$ and $q2$), and inversely proportional to the square of the distance, r, between them. (Potential energy is proportional to $q1q2/r$.) If the two charges are of opposite sign, the force between them is attractive; if they are of the same sign, the force is repulsive.

b. The first ionization energy is the minimum energy needed to remove the least tightly held electron from an atom or ion. In general, the ionization energy of any electron in an atom or ion is the minimum energy needed to remove that electron from the atom or ion.

c. The relative magnitude of the ionization energy can be estimated through qualitative application of Coulomb's Law. The farther an electron is from the nucleus, the lower its ionization energy. When comparing two species with the same arrangement of electrons, the higher the nuclear charge, the higher the ionization energy of an electron in a given subshell.

d. Photoelectron spectroscopy (PES) provides a useful means to engage students in the use of quantum mechanics to interpret spectroscopic data and extract information on atomic structure from such data. In particular, low-resolution PES of atoms provides direct evidence for the shell model. Light consists of photons, each of which has energy $E = h\nu$, where h is Planck's constant and ν is the frequency of the light. In the photoelectric effect, incident light ejects electrons from a material. This requires the photon to have sufficient energy to eject the electron. Photoelectron spectroscopy determines the energy needed to eject electrons from the material. Measurement of these energies provides a method to deduce the shell structure of an atom. The intensity of the photoelectron signal at a given energy is a measure of the number of electrons in that energy level.

e. The electronic structure of atoms with multiple electrons can be inferred from evidence provided by PES. For instance, both electrons in He are identical, and they are both roughly the same distance from the nucleus as in H, while there are two shells of electrons in Li, and the outermost electron is further from the nucleus than in H.

1.B.1 Learning Objectives:
LO 1.5 The student is able to explain the distribution of electrons in an atom or ion based upon data. [See SP 1.5, 6.2]

LO 1.6 The student is able to analyze data relating to electron energies for patterns and relationships. [See SP 5.1]

Essential knowledge 1.B.2: The electronic structure of the atom can be described using an electron configuration that reflects the concept of electrons in quantized energy levels or shells; the energetics of the electrons in the atom can be understood by consideration of Coulomb's Law.
a. Electron configurations provide a method for describing the distribution of electrons in an atom or ion.

b. Each electron in an atom has a different ionization energy, which can be qualitatively explained through Coulomb's Law.

c. In multielectron atoms and ions, the electrons can be thought of as being in "shells" and "subshells," as indicated by the relatively close ionization energies associated with some groups of electrons. Inner electrons are called core electrons, and outer electrons are called valence electrons.

d. Core electrons are generally closer to the nucleus than valence electrons, and they are considered to "shield" the valence electrons from the full electrostatic attraction of the nucleus. This phenomenon can be used in conjunction with Coulomb's Law to explain/rationalize/predict relative ionization energies. Differences in electron-electron repulsion are responsible for the differences in energy between electrons in different orbitals in the same shell.

1.B.2 Learning Objectives:
LO 1.7 The student is able to describe the electronic structure of the atom, using PES data, ionization energy data, and/or Coulomb's Law to construct explanations of how the energies of electrons within shells in atoms vary. [See SP 5.1, 6.2]

LO 1.8 The student is able to explain the distribution of electrons using Coulomb's Law to analyze measured energies. [See SP 6.2]

Enduring understanding 1.C: Elements display periodicity in their properties when the elements are organized according to increasing atomic number. This periodicity can be explained by the regular variations that occur in the electronic structures of atoms. Periodicity is a useful principle for understanding properties and predicting trends in properties. Its modern-day uses range from examining the composition of materials to generating ideas for designing new materials. Essential knowledge 1.C.1: Many properties of atoms exhibit periodic trends that are reflective of the periodicity of electronic structure.
a. The structure of the periodic table is a consequence of the pattern of electron configurations and the presence of shells (and subshells) of electrons in atoms.

b. Ignoring the few exceptions, the electron configuration for an atom can be deduced from the element's position in the periodic table.

c. For many atomic properties, trends within the periodic table (and relative values for different atoms and ions) can be qualitatively understood and explained using Coulomb's Law, the shell model, and the concept of shielding/effective nuclear charge. These properties include: 1. First ionization energy 2. Atomic and ionic radii 3. Electronegativity 4. Typical ionic charges

d. Periodicity is a useful tool when designing new molecules or materials, since replacing an element of one group with another of the same group may lead to a new substance with similar properties. For instance, since SiO2 can be a ceramic, SnO2 may be as well.

1.C.1 Learning Objectives:
LO 1.9 The student is able to predict and/or justify trends in atomic properties based on location on the periodic table and/or the shell model. [See SP 6.4]

LO 1.10 Students can justify with evidence the arrangement of the periodic table and can apply periodic properties to chemical reactivity. [See SP 6.1]

LO 1.11 The student can analyze data, based on periodicity and the properties of binary compounds, to identify patterns and generate hypotheses related to the molecular design of compounds for which data are not supplied. [See SP 3.1, 5.1]

Essential knowledge 1.C.2: The currently accepted best model of the atom is based on the quantum mechanical model.
a. Coulomb's Law is the basis for describing the energy of interaction between protons and electrons.

b. Electrons are not considered to follow specific orbits. Chemists refer to the region of space in which an electron is found as an orbital.

c. Electrons in atoms have an intrinsic property known as spin that can result in atoms having a magnetic moment. There can be at most two electrons in any orbital, and these electrons must have opposite spin.

d. The quantum mechanical (QM) model addresses known problems with the classical shell model and is also consistent with atomic electronic structures that correspond with the periodic table.

e. The QM model can be approximately solved using computers and serves as the basis for software that calculates the structure and reactivity of molecules.

1.C.2 Learning Objective:
LO 1.12 The student is able to explain why a given set of data suggests, or does not suggest, the need to refine the atomic model from a classical shell model with the quantum mechanical model. [See SP 6.3]

Enduring understanding 1.D: Atoms are so small that they are difficult to study directly; atomic models are constructed to explain experimental data on collections of atoms. Essential knowledge 1.D.1: As is the case with all scientific models, any model of the atom is subject to refinement and change in response to new experimental results. In that sense, an atomic model is not regarded as an exact description of the atom, but rather a theoretical construct that fits a set of experimental data.
a. Scientists use experimental results to test scientific models. When experimental results are not consistent with the predictions of a scientific model, the model must be revised or replaced with a new model that is able to predict/explain the new experimental results. A robust scientific model is one that can be used to explain/predict numerous results over a wide range of experimental circumstances.

b. The construction of a shell model of the atom through ionization energy information provides an opportunity to show how a model can be refined and changed as additional information is considered.

1.D.1 Learning Objective:
LO 1.13 Given information about a particular model of the atom, the student is able to determine if the model is consistent with specified evidence. [See SP 5.3]

Essential knowledge 1.D.3: The interaction of electromagnetic waves or light with matter is a powerful means to probe the structure of atoms and molecules, and to measure their concentration.
a. The energy of a photon is related to the frequency of the electromagnetic wave through Planck's equation ($E = h\nu$). When a photon is absorbed (or emitted) by a molecule, the energy of the molecule is increased (or decreased) by an amount equal to the energy of the photon.

b. Different types of molecular motion lead to absorption or emission of photons in different spectral regions. Infrared radiation is associated with transitions in molecular vibrations and so can be used to detect the presence of different types of bonds. Ultraviolet/visible radiation is associated with transitions in electronic energy levels and so can be used to probe electronic structure.

c. The amount of light absorbed by a solution can be used to determine the concentration of the absorbing molecules in that solution, via the Beer-Lambert Law.

1.D.3 Learning Objectives:
LO 1.15 The student can justify the selection of a particular type of spectroscopy to measure properties associated with vibrational or electronic motions of molecules. [See SP 4.1]

Big Idea 2: Chemical and physical properties of materials can be explained by the structure and the arrangement of atoms, ions, or molecules and the forces between them.
2.A Learning Objectives:
LO 2.1 Students can predict properties of substances based on their chemical formulas, and provide explanations of their properties based on particle views. [See SP 6.4, 7.1]

Essential knowledge 2.A.1: The different properties of solids and liquids can be explained by differences in their structures, both at the particulate level and in their supramolecular structures.
a. Solids can be crystalline, where the particles are arranged in a regular 3-D structure, or they can be

amorphous, where the particles do not have a regular, orderly arrangement. In both cases, the motion of the individual particles is limited, and the particles do not undergo any overall translation with respect to each other. Interparticle interactions and the ability to pack the particles together provide the main criteria for the structures of solids.

b. The constituent particles in liquids are very close to each other, and they are continually moving and colliding. The particles are able to undergo translation with respect to each other and their arrangement, and movement is influenced by the nature and strength of the intermolecular forces that are present.

c. The solid and liquid phases for a particular substance generally have relatively small differences in molar volume because in both cases the constituent particles are very close to each other at all times.

d. The differences in other properties, such as viscosity, surface tension, and volumes of mixing (for liquids), and hardness and macroscopic crystal structure (for solids), can be explained by differences in the strength of attraction between the particles and/or their overall organization.

e. Heating and cooling curves for pure substances provide insight into the energetics of liquid/solid phase changes.

2.A.1 Learning Objective:
LO 2.3 The student is able to use aspects of particulate models (i.e., particle spacing, motion, and forces of attraction) to reason about observed differences between solid and liquid phases and among solid and liquid materials. [See SP 6.4, 7.1]

Essential knowledge 2.A.2: The gaseous state can be effectively modeled with a mathematical equation relating various macroscopic properties. A gas has neither a definite volume nor a definite shape; because the effects of attractive forces are minimal, we usually assume that the particles move independently.
e. Some real gases exhibit ideal or near-ideal behavior under typical laboratory conditions. Laboratory data can be used to generate or investigate the relationships in 2.A.2.a and to estimate absolute zero on the Celsius scale.

f. All real gases are observed to deviate from ideal behavior, particularly under conditions that are close to those resulting in condensation. Except at extremely high pressures that are not typically seen in the laboratory, deviations from ideal behavior are the result of intermolecular attractions among gas molecules. These forces are strongly distance-dependent, so they are most significant during collisions.

g. Observed deviations from ideal gas behavior can be explained through an understanding of the structure of atoms and molecules and their intermolecular interactions.

Essential knowledge 2.A.3: Solutions are homogenous mixtures in which the physical properties are dependent on the concentration of the solute and the strengths of all interactions among the particles of the solutes and solvent.
a. In a solution (homogeneous mixture), the macroscopic properties do not vary throughout the sample. This is in contrast to a heterogeneous mixture in which the macroscopic properties depend upon the location in the mixture. The distinction between heterogeneous and homogeneous depends on the length scale of interest. As an example, colloids may be heterogeneous on the scale of micrometers, but homogeneous on the scale of centimeters.

b. Solutions come in the form of solids, liquids, and gases.

c. For liquid solutions, the solute may be a gas, a liquid, or a solid.

d. Based on the reflections of their structure on the microscopic scale, liquid solutions exhibit several general properties: 1. The components cannot be separated by using filter paper. 2. There are no components large enough to scatter visible light. 3. The components can be separated using processes that are a result of the intermolecular interactions between and among the components.

e. Chromatography (paper and column) separates chemical species by taking advantage of the differential strength of intermolecular interactions between and among the components.

f. Distillation is used to separate chemical species by taking advantage of the differential strength of intermolecular interactions between and among the components and the effects these interactions have on the vapor pressures of the components in the mixture.

g. The formation of a solution may be an exothermic or endothermic process, depending on the relative strengths of intermolecular/interparticle interactions before and after the dissolution process.

h. Generally, when ionic compounds are dissolved in water, the component ions are separated and dispersed. The presence of ions in a solution can be detected by use of conductivity measurements.

i. Solution composition can be expressed in a variety of ways; molarity is the most common method used in the laboratory. Molarity is defined as the number of moles of solute per liter of solution.

j. Understanding how to prepare solutions of specified molarity through direct mixing of the components, through use of volumetric glassware, and by dilution of a solution of known molarity with additional solvent is important for performing laboratory work in chemistry.

2.A.3 Learning Objectives:
LO 2.7 The student is able to explain how solutes can be separated by chromatography based on intermolecular interactions. [See SP 6.2]

LO 2.8 The student can draw and/or interpret representations of solutions that show the interactions between the solute and solvent. [See SP 1.1, 1.2, 6.4]

LO 2.9 The student is able to create or interpret representations that link the concept of molarity with particle views of solutions. [See SP 1.1, 1.4]

LO 2.10 The student can design and/or interpret the results of a separation experiment (filtration, paper chromatography, column chromatography, or distillation) in terms of the relative strength of interactions among and between the components. [See SP 4.2, 5.1]

Enduring understanding 2.B: Forces of attraction between particles (including the noble gases and also different parts of some large molecules) are important in determining many macroscopic properties of a substance, including how the observable physical state changes with temperature. Essential knowledge 2.B.1: London dispersion forces are attractive forces present between all atoms and molecules. London dispersion forces are often the strongest net intermolecular force between large molecules.
a. A temporary, instantaneous dipole may be created by an uneven distribution of electrons around the nucleus (nuclei) of an atom (molecule).

b. London dispersion forces arise due to the Coulombic interaction of the temporary dipole with the electron distribution in neighboring atoms and molecules.

c. Dispersion forces increase with contact area between molecules and with increasing polarizability of the molecules. The polarizability of a molecule increases with the number of electrons in the molecule, and is enhanced by the presence of pi bonding.

2.B.1 Learning Objective:

LO 2.11 The student is able to explain the trends in properties and/or predict properties of samples consisting of particles with no permanent dipole on the basis of London dispersion forces. [See SP 6.2, 6.4]

Essential knowledge 2.B.2: Dipole forces result from the attraction among the positive ends and negative ends of polar molecules. Hydrogen bonding is a strong type of dipole-dipole force.
a. Molecules with dipole moments experience Coulombic interactions that result in a net attractive interaction when they are near each other. 1. Intermolecular dipole-dipole forces are weaker than ionic forces or covalent bonds. 2. Interactions between polar molecules are typically greater than between nonpolar molecules of comparable size because these interactions act in addition to London dispersion forces. 3. Dipole-dipole attractions can be represented by diagrams of attraction between the positive and negative ends of polar molecules trying to maximize attractions and minimize repulsions in the liquid or solid state. 4. Dipole-induced dipole interactions are present between a polar and nonpolar molecule. The strength of these forces increases with the magnitude of the dipole of the polar molecule and with the polarizability of the nonpolar molecule.

b. Hydrogen bonding is a relatively strong type of intermolecular interaction that occurs when hydrogen atoms that are covalently bonded to the highly electronegative atoms (N, O, and F) are also attracted to the negative end of a dipole formed by the electronegative atom (N, O, and F) in a different molecule, or a different part of the same molecule. When hydrogen bonding is present, even small molecules may have strong intermolecular attractions.

c. Hydrogen bonding between molecules, or between different parts of a single molecule, may be represented by diagrams of molecules with hydrogen bonding and indications of location of hydrogen bonding.

d. Ionic interactions with dipoles are important in the solubility of ionic compounds in polar solvents.

2.B.2 Learning Objectives:
LO 2.13 The student is able to describe the relationships between the structural features of polar molecules and the forces of attraction between the particles. [See SP 1.4]

LO 2.14 The student is able to apply Coulomb’s Law qualitatively (including using representations) to describe the interactions of ions, and the attractions between ions and solvents to explain the factors that contribute to the solubility of ionic compounds. [See SP 1.4, 6.4]

Essential knowledge 2.B.3: Intermolecular forces play a key role in determining the properties of substances, including biological structures and interactions.
a. Many properties of liquids and solids are determined by the strengths and types of intermolecular forces present. 1. Boiling point 2. Surface tension 3. Capillary action 4. Vapor pressure

b. Substances with similar intermolecular interactions tend to be miscible or soluble in one another.

c. The presence of intermolecular forces among gaseous particles, including noble gases, leads to deviations from ideal behavior, and it can lead to condensation at sufficiently low temperatures and/or sufficiently high pressures.

d. Graphs of the pressure-volume relationship for real gases can demonstrate the deviation from ideal behavior; these deviations can be interpreted in terms of the presence and strengths of intermolecular forces.

e. The structure and function of many biological systems depend on the strength and nature of the various Coulombic forces. 1. Substrate interactions with the active sites in enzyme catalysis 2. Hydrophilic and hydrophobic regions in proteins that determine three-dimensional structure in water solutions

2.B.3 Learning Objectives:

LO 2.15 The student is able to explain observations regarding the solubility of ionic solids and molecules in water and other solvents on the basis of particle views that include intermolecular interactions and entropic effects. [See SP 1.4, 6.2, connects to 5.E.1]

LO 2.16 The student is able to explain the properties (phase, vapor pressure, viscosity, etc.) of small and large molecular compounds in terms of the strengths and types of intermolecular forces. [See SP 6.2]

Enduring understanding 2.C: The strong electrostatic forces of attraction holding atoms together in a unit are called chemical bonds.
a. The strong electrostatic forces of attraction holding atoms together in a unit are called chemical bonds.

2.C Learning Objective:
LO 2.17 The student can predict the type of bonding present between two atoms in a binary compound based on position in the periodic table and the electronegativity of the elements. [See SP 6.4]

Essential knowledge 2.C.1: In covalent bonding, electrons are shared between the nuclei of two atoms to form a molecule or polyatomic ion. Electronegativity differences between the two atoms account for the distribution of the shared electrons and the polarity of the bond.
a. Electronegativity is the ability of an atom in a molecule to attract shared electrons to it.

b. Electronegativity values for the representative elements increase going from left to right across a period and decrease going down a group. These trends can be understood qualitatively through the electronic structure of the atoms, the shell model, and Coulomb's Law.

c. Two or more valence electrons shared between atoms of identical electronegativity constitute a nonpolar covalent bond.

d. However, bonds between carbon and hydrogen are often considered to be nonpolar even though carbon is slightly more electronegative than hydrogen. The formation of a nonpolar covalent bond can be represented graphically as a plot of potential energy vs. distance for the interaction of two identical atoms. Hydrogen atoms are often used as an example. 1. The relative strengths of attractive and repulsive forces as a function of distance determine the shape of the graph. 2. The bond length is the distance between the bonded atoms' nuclei, and is the distance of minimum potential energy where the attractive and repulsive forces are balanced. 3. The bond energy is the energy required for the dissociation of the bond. This is the net energy of stabilization of the bond compared to the two separated atoms. Typically, bond energy is given on a per mole basis.

e. Two or more valence electrons shared between atoms of unequal electronegativity constitute a polar covalent bond. 1. The difference in electronegativity for the two atoms involved in a polar covalent bond is not equal to zero. 2. The atom with a higher electronegativity will develop a partial negative charge relative to the other atom in the bond. For diatomic molecules, the partial negative charge on the more electronegative atom is equal in magnitude to the partial positive charge on the less electronegative atom. 3. Greater differences in electronegativity lead to greater partial charges, and consequently greater bond dipoles. 4. The sum of partial charges in any molecule or ion must be equal to the overall charge on the species.

f. All bonds have some ionic character, and the difference between ionic and covalent bonding is not distinct but rather a continuum. The difference in electronegativity is not the only factor in determining if a bond is designated ionic or covalent. Generally, bonds between a metal and nonmetal are ionic, and between two nonmetals the bonds are covalent. Examination of the properties of a compound is the best way to determine the type of bonding.

2.C.1 Learning Objective:
LO 2.18 The student is able to rank and justify the ranking of bond polarity on the basis of the locations of the

bonded atoms in the periodic table. [See SP 6.1]

Essential knowledge 2.C.2: Ionic bonding results from the net attraction between oppositely charged ions, closely packed together in a crystal lattice.
a. The cations and anions in an ionic crystal are arranged in a systematic, periodic 3-D array that maximizes the attractive forces among cations and anions while minimizing the repulsive forces.

b. Coulomb's Law describes the force of attraction between the cations and anions in an ionic crystal. 1. Because the force is proportional to the charge on each ion, larger charges lead to stronger interactions. 2. Because the force is inversely proportional to the square of the distance between the centers of the ions (nuclei), smaller ions lead to stronger interactions.

2.C.2 Learning Objective:
LO 2.19 The student can create visual representations of ionic substances that connect the microscopic structure to macroscopic properties, and/or use representations to connect the microscopic structure to macroscopic properties (e.g., boiling point, solubility, hardness, brittleness, low volatility, lack of malleability, ductility, or conductivity). [See SP 1.1, 1.4, 7.1, connects to 2.D.1, 2.D.2]

Essential knowledge 2.C.3: Metallic bonding describes an array of positively charged metal cores surrounded by a sea of mobile valence electrons.
a. The valence electrons from the metal atoms are considered to be delocalized and not associated with any individual atom.

b. Metallic bonding can be represented as an array of positive metal ions with valence electrons drawn among them, as if the electrons were moving (i.e., a sea of electrons).

c. The electron sea model can be used to explain several properties of metals, including electrical conductivity, malleability, ductility, and low volatility.

d. The number of valence electrons involved in metallic bonding, via the shell model, can be used to understand patterns in these properties, and can be related to the shell model to reinforce the connections between metallic bonding and other forms of bonding.

2.C.3 Learning Objective:
LO 2.20 The student is able to explain how a bonding model involving delocalized electrons is consistent with macroscopic properties of metals (e.g., conductivity, malleability, ductility, and low volatility) and the shell model of the atom. [See SP 6.2, 7.1, connects to 2.D.2]

Essential knowledge 2.C.4: The localized electron bonding model describes and predicts molecular geometry using Lewis diagrams and the VSEPR model.
a. Lewis diagrams can be constructed according to a well-established set of principles.

b. The VSEPR model uses the Coulombic repulsion between electrons as a basis for predicting the arrangement of electron pairs around a central atom.

c. In cases where more than one equivalent Lewis structure can be constructed, resonance must be included as a refinement to the Lewis structure approach in order to provide qualitatively accurate predictions of molecular structure and properties (in some cases).

d. Formal charge can be used as a criterion for determining which of several possible valid Lewis diagrams provides the best model for predicting molecular structure and properties.

e. The combination of Lewis diagrams with the VSEPR model provides a powerful model for predicting structural properties of many covalently bonded molecules and polyatomic ions, including the following: 1. Molecular geometry 2. Bond angles 3. Relative bond energies based on bond order 4. Relative bond lengths (multiple bonds, effects of atomic radius) 5. Presence of a dipole moment

f. As with any model, there are limitations to the use of the Lewis structure model, particularly in cases with an odd number of valence electrons. Recognizing that Lewis diagrams have limitations is of significance.

g. Organic chemists commonly use the terms "hybridization" and "hybrid orbital" to describe the arrangement of electrons around a central atom. When there is a bond angle of 180°, the central atom is said to be sp hybridized; for 120°, the central atom is sp2 hybridized; and for 109°, the central atom is sp3 hybridized. Students should be aware of this terminology, and be able to use it. When an atom has more than four pairs of electrons surrounding the central atom, students are only responsible for the shape of the resulting molecule.

h. Bond formation is associated with overlap between atomic orbitals. In multiple bonds, such overlap leads to the formation of both sigma and pi bonds. The overlap is stronger in sigma than pi bonds, which is reflected in sigma bonds having larger bond energy than pi bonds. The presence of a pi bond also prevents the rotation of the bond, and leads to structural isomers. In systems, such as benzene, where atomic p-orbitals overlap strongly with more than one other p-orbital, extended pi bonding exists, which is delocalized across more than two nuclei. Such descriptions provide an alternative description to resonance in Lewis structures. A useful example of delocalized pi bonding is molecular solids that conduct electricity. The discovery of such materials at the end of the 1970s overturned a long-standing assumption in chemistry that molecular solids will always be insulators.

i. Molecular orbital theory describes covalent bonding in a manner that can capture a wider array of systems and phenomena than the Lewis or VSEPR models. Molecular orbital diagrams, showing the correlation between atomic and molecular orbitals, are a useful qualitative tool related to molecular orbital theory.

2.C.4 Learning Objective:
LO 2.21 The student is able to use Lewis diagrams and VSEPR to predict the geometry of molecules, identify hybridization, and make predictions about polarity. [See SP 1.4]

Enduring understanding 2.D: The type of bonding in the solid state can be deduced from the properties of the solid state.
a. The type of bonding in the solid state can be deduced from the properties of the solid state.

2.D Learning Objective:
LO 2.22 The student is able to design or evaluate a plan to collect and/or interpret data needed to deduce the type of bonding in a sample of a solid. [See SP 4.2]

Essential knowledge 2.D.1: Ionic solids have high melting points, are brittle, and conduct electricity only when molten or in solution.
a. Many properties of ionic solids are related to their structure. 1. Ionic solids generally have low vapor pressure due to the strong Coulombic interactions of positive and negative ions arranged in a regular three-dimensional array. 2. Ionic solids tend to be brittle due to the repulsion of like charges caused when one layer slides across another layer. 3. Ionic solids do not conduct electricity. However, when ionic solids are melted, they do conduct electricity because the ions are free to move. 4. When ionic solids are dissolved in water, the separated ions are free to move; therefore, these solutions will conduct electricity. Dissolving a nonconducting solid in water, and observing the solution's ability to conduct electricity, is one way to identify an ionic solid. 5. Ionic compounds tend not to dissolve in nonpolar solvents because the attractions among the ions are much stronger than the attractions among the separated ions and the nonpolar solvent molecules.

b. The attractive force between any two ions is governed by Coulomb's Law: The force is directly proportional to the charge of each ion and inversely proportional to the square of the distance between the centers of the ions. 1. For ions of a given charge, the smaller the ions, and thus the smaller the distance between ion centers, the stronger the Coulombic force of attraction, and the higher the melting point. 2. Ions with higher charges lead to higher Coulombic forces, and therefore higher melting points.

2.D.1 Learning Objectives:
LO 2.23 The student can create a representation of an ionic solid that shows essential characteristics of the structure and interactions present in the substance. [See SP 1.1]

LO 2.24 The student is able to explain a representation that connects properties of an ionic solid to its structural attributes and to the interactions present at the atomic level. [See SP 1.1, 6.2, 7.1]

Essential knowledge 2.D.2: Metallic solids are good conductors of heat and electricity, have a wide range of melting points, and are shiny, malleable, ductile, and readily alloyed.
a. A metallic solid can be represented as positive kernels (or cores) consisting of the nucleus and inner electrons of each atom surrounded by a sea of mobile valence electrons. 1. Metals are good conductors because the electrons are delocalized and relatively free to move. 2. Metals are malleable and ductile because deforming the solid does not change the environment immediately surrounding each metal core.

b. Metallic solids are often pure substances, but may also be mixtures called alloys. 1. Some properties of alloys can be understood in terms of the size of the component atoms: — Interstitial alloys form between atoms of different radius, where the smaller atoms fill the interstitial spaces between the larger atoms. (Steel is an example in which carbon occupies the interstices in iron.) The interstitial atoms do not appreciably expand the lattice, so the density is often substantially increased. The interstitial atoms make the lattice more rigid, decreasing malleability and ductility. — Substitutional alloys form between atoms of comparable radius, where one atom substitutes for the other in the lattice. (Brass is an example in which some copper atoms are substituted with a different element, usually zinc.) The density typically lies between those of the component metals, and the alloy remains malleable and ductile. 2. Alloys typically retain a sea of mobile electrons and so remain conducting. 3. In some cases, alloy formation alters the chemistry of the surface. An example is formation of a chemically inert oxide layer in stainless steel.

2.D.2 Learning Objectives:
LO 2.25 The student is able to compare the properties of metal alloys with their constituent elements to determine if an alloy has formed, identify the type of alloy formed, and explain the differences in properties using particulate level reasoning. [See SP 1.4, 7.2]

LO 2.26 Students can use the electron sea model of metallic bonding to predict or make claims about the macroscopic properties of metals or alloys. [See SP 6.4, 7.1]

LO 2.27 The student can create a representation of a metallic solid that shows essential characteristics of the structure and interactions present in the substance. [See SP 1.1]

LO 2.28 The student is able to explain a representation that connects properties of a metallic solid to its structural attributes and to the interactions present at the atomic level. [See SP 1.1, 6.2, 7.1]

Essential knowledge 2.D.3: Covalent network solids generally have extremely high melting points, are hard, and are thermal insulators. Some conduct electricity.
a. Covalent network solids consist of atoms that are covalently bonded together into a two-dimensional or three-dimensional network. 1. Covalent network solids are only formed from nonmetals: elemental (diamond, graphite) or two nonmetals (silicon dioxide and silicon carbide). 2. The properties of covalent network solids are a reflection of their structure. 3. Covalent network solids have high melting points because all of the atoms are covalently bonded. 4. Three-dimensional covalent networks tend to be rigid and hard because the covalent bond angles are fixed. 5. Generally, covalent network solids form in the carbon group because of their ability

to form four covalent bonds.

b. Graphite is an allotrope of carbon that forms sheets of two-dimensional networks. 1. Graphite has a high melting point because the covalent bonds between the carbon atoms making up each layer are relatively strong. 2. Graphite is soft because adjacent layers can slide past each other relatively easily; the major forces of attraction between the layers are London dispersion forces.

c. Silicon is a covalent network solid and a semiconductor. 1. Silicon forms a three-dimensional network similar in geometry to a diamond. 2. Silicon's conductivity increases as temperature increases. 3. Periodicity can be used to understand why doping with an element with one extra valence electron converts silicon into an n-type semiconducting (negative charge carrying) material, while doping with an element with one less valence electron converts silicon into a p-type semiconducting (positive charge carrying) material. Junctions between n-doped and p-doped materials can be used to control electron flow, and thereby are the basis of modern electronics.

2.D.3 Learning Objectives:
LO 2.29 The student can create a representation of a covalent solid that shows essential characteristics of the structure and interactions present in the substance. [See SP 1.1]

LO 2.30 The student is able to explain a representation that connects properties of a covalent solid to its structural attributes and to the interactions present at the atomic level. [See SP 1.1, 6.2, 7.1]

Essential knowledge 2.D.4: Molecular solids with low molecular weight usually have low melting points and are not expected to conduct electricity as solids, in solution, or when molten.
a. Molecular solids consist of nonmetals, diatomic elements, or compounds formed from two or more nonmetals.

b. Molecular solids are composed of distinct, individual units of covalently bonded molecules attracted to each other through relatively weak intermolecular forces. 1. Molecular solids are not expected to conduct electricity because their electrons are tightly held within the covalent bonds of each constituent molecule. 2. Molecular solids generally have a low melting point because of the relatively weak intermolecular forces present between the molecules. 3. Molecular solids are sometimes composed of very large molecules, or polymers, with important commercial and biological applications.

2.D.4 Learning Objectives:
LO 2.31 The student can create a representation of a molecular solid that shows essential characteristics of the structure and interactions present in the substance. [See SP 1.1]

LO 2.32 The student is able to explain a representation that connects properties of a molecular solid to its structural attributes and to the interactions present at the atomic level. [See SP 1.1, 6.2, 7.1]

Big Idea 5: The laws of thermodynamics describe the essential role of energy and explain and predict the direction of changes in matter.
5.A Learning Objective:
LO 5.1 The student is able to create or use graphical representations in order to connect the dependence of potential energy to the distance between atoms and factors, such as bond order (for covalent interactions) and polarity (for intermolecular interactions), which influence the interaction strength. [See SP 1.1, 1.4, 7.2, connects to Big Idea 2] Note: This learning objective applies to essential knowledge components of 5A–5E.

Enduring understanding 5.C: Breaking bonds requires energy, and making bonds releases energy. Essential knowledge 5.C.1: Potential energy is associated with a particular geometric arrangement of atoms or ions and the electrostatic interactions between them.
a. The attraction between the electrons of one atom and the protons of another explains the tendency for the

atoms to approach one another. The repulsion between the nuclei (or core electrons) explains why the atoms repel one another at close distance. The distance at which the energy of interaction is minimized is called the bond length, and the atoms vibrate about this minimum energy position.

b. A graph of energy versus the distance between atoms can be plotted and interpreted. Using this graph, it is possible to identify bond length and bond energy.

c. Conceptually, bond making and bond breaking are opposing processes that have the same magnitude of energy associated with them. Thus, convention becomes important, so we define the bond energy as the energy required to break a bond.

d. Because chemical bonding arises from electrostatic interaction between electrons and nuclei, larger charges tend to lead to larger strengths of interaction. Thus, triple bonds are stronger than double or single bonds because they share more pairs of electrons.

e. Stronger bonds tend to be shorter bonds.

Essential knowledge 5.C.2: The net energy change during a reaction is the sum of the energy required to break the bonds in the reactant molecules and the energy released in forming the bonds of the product molecules. The net change in energy may be positive for endothermic reactions where energy is required, or negative for exothermic reactions where energy is released.
a. During a chemical reaction, bonds are broken and/or formed, and these events change the potential energy of the reaction system.

b. The average energy required to break all of the bonds in the reactant molecules can be estimated by adding up the average bond energies or bond enthalpies for all the bonds in the reactant molecules. Likewise, the average energy released in forming the bonds in the products can be estimated. If the energy released is greater than the energy required, then the reaction is exothermic. If the energy required is greater than the energy released, then the reaction is endothermic.

c. For an exothermic reaction, the products are at a lower potential energy compared with the reactants. For an endothermic reaction, the products are at a higher potential energy than the reactants.

5.C.2 Learning Objective:
LO 5.8 The student is able to draw qualitative and quantitative connections between the reaction enthalpy and the energies involved in the breaking and formation of chemical bonds. [See SP 2.3, 7.1, 7.2]

Big Idea 6: Any bond or intermolecular attraction that can be formed can be broken. These two processes are in a dynamic competition, sensitive to initial conditions and external perturbations.
Enduring understanding 6.A: Chemical equilibrium is a dynamic, reversible state in which rates of opposing processes are equal. Essential knowledge 6.A.1: In many classes of reactions, it is important to consider both the forward and reverse reaction.
a. Many readily observable processes are reversible. Examples include evaporating and condensing water, absorption of a gas, or dissolving and precipitating a salt. Relevant and interesting contexts include biological examples (binding of oxygen to hemoglobin and the attachment of molecules to receptor sites in the nose) and environmental examples (transfer of carbon between atmosphere and biosphere and transfer of dissolved substances between atmosphere and hydrosphere).

b. Dissolution of a solid, transfer of protons in acid-base reactions, and transfer of electrons in redox reactions are important examples of reversible reactions.

6.A.1 Learning Objective:
LO 6.1 The student is able to, given a set of experimental observations regarding physical, chemical,

biological, or environmental processes that are reversible, construct an explanation that connects the observations to the reversibility of the underlying chemical reactions or processes. [See SP 6.2]

6.C.3 Learning Objectives:
LO 6.24 The student can analyze the enthalpic and entropic changes associated with the dissolution of a salt, using particulate level interactions and representations. [See SP 1.4, 7.1, connects to 5.E]