

Unit 1: Atomic Structure and Properties

Essential Questions:

- How can the same element be used in nuclear fuel rods and fake diamonds?
- How can large quantities of objects be counted by weighing?
- If atoms are too small to be observed directly, how do we know how they're structured?
- Why does the periodic table have the shape that it does?

Unit 1	Lesson Time	Suggested Skill	Learning Objective	Essential Knowledge
Topic 1.1 Moles and Molar Mass	40 minutes	5.B	LO 1.1.A Calculate quantities of a substance or its relative number of particles using dimensional analysis and the mole concept.	EK 1.1.A.1 EK 1.1.A.2 EK 1.1.A.3

Materials	Instructional Activities and Assessments
<i>Chemistry: An Atoms-Focused Approach</i> Chapter 2, 2.4–2.5, pp. 62–71 Connections Activity: two 100 mL beakers, table salt, aluminum foil, digital scale, water Extended practice: Smartwork questions: mass to moles or mass to particles Helpful Videos The Mole Mole Conversions The Mole and Avogadro's Number	Activate Prior Knowledge In order to understand the quantitative connection between Avogadro's number and a compound or element's mass, students must first be able to identify the types of particles to which Avogadro's number can be associated and where the mass of those particles can be found or calculated. Begin class by having students identify the mass of an element (choose something that is not diatomic; aluminum is recommended to connect to the activity below) and calculate the mass of a molecular and ionic compound. Familiar compounds like water for the molecular compound and sodium chloride for the ionic compound may help students make more immediate connections. Make Connections to New Learning Utilize the description below as a model for a conversation to allow students to connect the idea of molar mass or volume to Avogadro's number. Begin the lesson by prompting students to locate the atomic mass of aluminum on the periodic table. Confirm that it is 26.98 grams, and then invite a student to measure out that amount using a scale, guiding them to use proper measurement techniques. Set the sample aside. Next, ask the class to share the calculated mass of sodium chloride. When a student responds with 58.44 grams, affirm the answer and ask them to explain their process. Reinforce that the formula mass of a compound is determined by adding the atomic masses of each element from the periodic table. Invite another student to measure out that mass of sodium chloride into a beaker, again emphasizing accurate measurement. Repeat this process with water, confirming the molecular mass as 18.02 grams, and have a student measure that amount. Once all three samples are prepared, ask the class to consider what these samples have in common. Guide the discussion toward the concept that each sample contains the same number of particles—6.02×10^{23} atoms, molecules, or formula units—introducing Avogadro's number as the key to understanding how molar mass and volume relate to particle count in a substance.

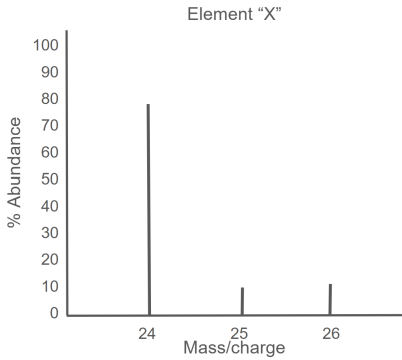
This is a preliminary sample; all content subject to change.

	Apply New Knowledge
--	----------------------------

	Utilize chapter problems 2.65 - 2.82 as practice for students during class as they work to convert mass to moles or particles. Guide the class through problem solving by modeling how to solve a problem using dimensional analysis. Have students solve subsequent practice problems in partners, checking progress as they work.
--	---

This is a preliminary sample; all content subject to change.

Unit 1	Lesson Time	Suggested Skill	Learning Objective	Essential Knowledge
Topic 1.2 Mass Spectra of Elements	20 minutes	5.D	LO 1.2.A Explain the quantitative relationship between the mass spectrum of an element and the masses of the element's isotopes.	EK 1.2.A.1 EK 1.2.A.2

Materials	Instructional Activities and Assessments
<i>Chemistry: An Atoms-Focused Approach</i> Chapter 2, 2.6, p. 75 Simulation “A Simplified Mass Spectrometer” Make Connections: Graph Image  Smartwork problem with tutorial 2.5 Chapter 2 practice questions 2.54–2.56 Helpful Video Mass Spectrometry of Elements	Activate Prior Knowledge Ask students to look at the periodic table and find elements with atomic masses that are whole numbers. Students most likely will note that elements near the bottom, starting with polonium, generally have whole number masses while those near the top have decimal values for masses. Share that decimal values indicate that those elements have alternate forms of their atoms called isotopes. Ask students to define isotope. They should know that isotopes are atoms of the same element that have varied masses due to differing amounts of neutrons. Make Connections to New Learning This concept revolves around students understanding the mass spectrometry graph. The goal is for students to connect the data in this graph to the average atomic mass found in the periodic table. A sample dialog is below to engage students with the graph, as it is often shown on the exam. Begin by drawing a graph on the board, labeling only the y-axis as "percent abundance", and ask students what they think that term means. Use their responses to guide a brief discussion, then explain that percent abundance refers to the proportion of each isotope of an element found in nature. Add a title to the graph, such as "Element X", to indicate that the data represents an unknown element. Ask students what varies between isotopes, leading them to identify neutron count, which is reflected in the mass. Label the x-axis with three isotope masses: 24, 25, and 26, and ask how many isotopes this element has. Once students identify three, draw bars to represent their relative abundances: approximately 78% for 24X, 10% for 25X, and 12% for 26X. Ask students which isotope is most abundant, confirming that 24X is the most common. Then, transition into a demonstration of how to calculate the average atomic mass using the data. Show how to convert percentages into decimals and multiply each by its respective mass: Average Atomic Mass = $(0.78 \times 24) + (0.10 \times 25) + (0.12 \times 26) = 24.34 \text{ g}$. Use this calculation to reinforce the concept that average atomic mass reflects both the mass and natural abundance of each isotope. Note: There are three types of problems involving average atomic mass: 1. Calculation of average atomic mass given percentages of isotopes 2. Calculation of the natural abundance of each isotope 3. Calculation of the exact mass of an isotope

This is a preliminary sample; all content subject to change.

The mass spectra on the AP® exam will be limited to atomic masses of single elements. The graph axes may be labeled differently, but the concept is the same as that discussed in the book.

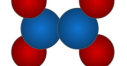
Apply New Knowledge

Utilize Smartwork problem 2.5 to review the tutorial on percent abundance with the class or have students navigate the problem by themselves while monitoring the room. For additional practice, utilize text questions as shown in the materials for an exit ticket.

If time allows, students can create their own mass spectrum of an element (drawn by hand) and allow another student to do the calculations to see if the mass spectrum is correct. The ChemTour video “Simple Mass Spectrometer” is a quick (~2 minutes) and helpful visual of how a mass spectrometer is used to determine isotopic composition.

This is a preliminary sample; all content subject to change.

Unit 1	Lesson Time	Suggested Skill	Learning Objective	Essential Knowledge
Topic 1.3 Elemental Composition of Pure Substances	25 minutes	2.A	LO 1.3.A Explain the quantitative relationship between the elemental composition by mass and the empirical formula of a pure substance.	EK 1.3.A.1 EK 1.3.A.2 EK 1.3.A.3

Materials	Instructional Activities and Assessments
<i>Chemistry: An Atoms-Focused Approach</i> Chapter 1, 1.5, pp. 16–17 Chapter 7, 7.5, pp. 317–320 Empirical Formula H₂O  HO  NO₂  Molecular Formula H₂O  H₂O₂  N₂O₄  You Tube: Finding and Calculating an Empirical Formula of a Compound Simulation “Percent Composition” Sections 7 and 8 Chapter 7 Practice Questions: 7.69–7.74	Activate Prior Knowledge As shown in a prior lesson, experimental data can be used to determine the composition of a sample or compound. In this lesson, we will explore how a formula called an empirical formula can be found using experimental data. Note that empirical formulas are introduced briefly in chapter 1 and in depth in chapter 7. Share the lesson learning outcome above with students to set the stage, then ask them for the formula for water. Ask them what would happen if, rather than a 2:1 ratio of hydrogen to oxygen was used, the result of using a 2:2 ratio of those elements? Allow them to discuss for 1-2 minutes. Make Connections to New Learning The Law of Definite Proportions states that a given chemical compound always contains its constituent elements in a fixed ratio by mass, regardless of its source or quantity. In contrast, the Law of Multiple Proportions applies when two elements form more than one compound; it asserts that the masses of one element that combine with a fixed mass of the other element will be in a ratio of small, whole numbers. As an example, provide students with images of some common compounds (like those shown in materials). Emphasize that the empirical and molecular formulas may be the same, as in the case of water, but the molecule formula may be a whole number multiple of the empirical formula. Data about the percent composition of a compound is how we find the empirical formula. Memory Trick! When working problems where the empirical formula is to be found from the percent composition, students can use the poem below to remember the steps to solve the problem. Percent to mass, Mass to mole, Divide by smallest, And round (or multiply) ‘till whole. Students can assume the percents given, if totalling 100%, can be used as the mass of each element in the compound. Divide the mass by the molar mass of the element to convert to moles. Divide each mole amount by the smallest to get a ratio, then round (if close to a whole number) or multiple to get whole numbers, which become the subscripts of the

This is a preliminary sample; all content subject to change.

	empirical formula. The YouTube video in the materials can provide further insight for students who are unsure how to apply the poem. Apply New Knowledge Students should practice solving for an empirical formula from percent composition before moving to the next lesson. Suggested practice problems are listed in the materials section. These problems can be worked in a scaffolded manner with the teacher leading the whole class or as group or individual work.
--	---

This is a preliminary sample; all content subject to change.

Unit 1	Lesson Time	Suggested Skill	Learning Objective	Essential Knowledge
Topic 1.4 Composition of Mixtures	30 minutes	5.A	LO 1.4.A Explain the quantitative relationship between the elemental composition by mass and the composition of substances in a mixture.	EK 1.4.A.1 EK 1.4.A.2

Materials	Instructional Activities and Assessments
Mixture: sand, salt or sugar, styrofoam beads or crumbled packing peanuts Paper plates One set of samples of pure aluminum, water, and sugar. <i>Chemistry: An Atoms-Focused Approach</i> Chapter 1, 1.3, pp. 10–13 Chapter 7, 7.5, pp. 317–319 Simulation “Percent Composition” Sections 1–6 Chapter 7 Practice Questions: 7.59–7.66 Chromatography Activity: Chromatography or filter paper, water soluble markers for each student, water, small beakers, tape Chromatography Experiment from Society of Women Engineers	Activate Prior Knowledge Provide groups of students with sealed containers of a mixture (see materials) and a paper plate. Allow them to pour the mixture onto the plate and identify the components. Ask students if what they are observing is a mixture or a pure substance. Ask them to justify their reasoning. Students should identify that this is a mixture and may justify it by saying that it can be separated through physical means or that it has inconsistent make up (the layers will often separate in the jars). To deepen conversation, ask students how they might separate this mixture. Answers might include adding water to allow the styrofoam to float and dissolving the sugar. They might also pick out or sift out the styrofoam. The sugar can be dissolved in water, filtered from the sand, and then dried for recovery. Show the class the samples of water, aluminum, and sugar. Ask them if these are mixtures, compounds, or elements. Ask what differentiates elements from compounds. Students should know that elements are the basic building blocks of matter found on the periodic table while compounds are created when two or more different elements are chemically bonded together. Utilize definitions from chapter 1 (see materials) as needed for this conversation. Most year 2 chemistry students should be familiar with these topics. Make Connections to New Learning Help students connect what they know about compounds to the concept of percent composition. Using a simple example like water or carbon dioxide, remind students that not only can we find the molar mass of the compound by adding the atomic masses of the atoms from which it is composed, we can also find the percent by mass that each element has in the compound. Calculate the percent mass of each element in a simple compound in front of the students, then work backwards from percentages to find the empirical formula (lesson tie in from AP CED topic 1.3). This will ensure students correctly connect that the percent composition can be solved for if the formula is known, or that the formula can be used to solve for percent composition. Apply New Knowledge

This is a preliminary sample; all content subject to change.

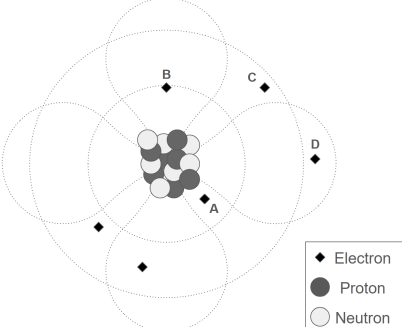
	As a class, in groups, or during independent practice, students should work examples of calculating the percent composition of a single element in a compound or find the percentages of all elements in a compound. See recommended problems in the materials.
--	--

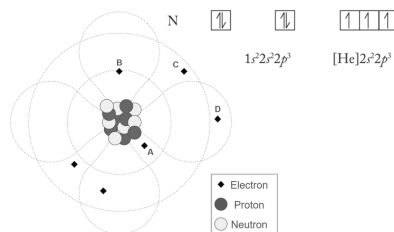
To reinforce the separation of mixtures via paper chromatography, perform the following:

Give each student a strip of chromatography paper or a strip of filter paper (a coffee filter will also work) and have them make a mark on the paper using a black, purple, or orange overhead marker (a water-soluble marker). Have them suspend the strip in a container of water, so that the water just touches the tip of the paper. At a convenient stopping point in class, have them pull the paper out, mark the end of the water front (also known as the solvent front), and dry the paper. As homework, have them describe what happened.

This is a preliminary sample; all content subject to change.

Unit 1	Lesson Time	Suggested Skill	Learning Objective	Essential Knowledge
Topic 1.5 Atomic Structure and Electron Configuration	75 minutes	1.A	LO 1.5.A Represent the ground-state electron configuration of an atom of an element or its ions using the Aufbau principle.	EK 1.5.A.1 EK 1.5.A.2 EK 1.5.A.3 EK 1.5.A.4

Materials	Instructional Activities and Assessments
<i>Chemistry: An Atoms-Focused Approach</i> Chapter 3, 3.4–3.9, pp. 97–123 Graphic for Activate Prior Knowledge  YouTube: Writing the Electron Configuration of an Atom Using the Periodic Table Orbital Filling Diagram, p. 116 Diagram with Orbital Filling Example	Activate Prior Knowledge For this lesson, students should have a solid foundation of vocabulary to include: nucleus, proton, neutron, electron, orbitals (s, p, d, f), attraction, repulsion, neutral, ion. It is important for students to know the concept that, when electrons absorb energy, they move to a higher energy state and then return to a lower energy state by emitting energy (light). Begin class with a quick activity. Provide a diagram like the one shown in the materials to partner pairs of students. Place the key vocabulary on the board and ask students to try to use all words as they write 4-5 sentences describing everything they know about the picture. Note: Scaffold the activity by allowing students to use terms as simple labels or to create a word web rather than a paragraph to relate terms. <i>Example student answer: The center of this image is the nucleus and it is made of protons and neutrons. Electrons are the diamonds and they are outside the nucleus in orbitals. The electrons are attracted to the protons but are repulsed by one another because like charges experience repulsion and opposite charges attract. The circular orbitals are s orbitals. The figure-8 orbitals are p orbitals. The d and f orbitals are not shown because this is a neutral atom of nitrogen and it does not have d or f orbitals. An ion of nitrogen would have more or less electrons than protons.</i> Have students share their use of the key vocabulary, correcting misconceptions as needed before beginning the lesson. Make Connections to New Learning The final learning output of this lesson is for students to be able to write ground state electron configurations for neutral atoms or ions. New learning for most second year students should focus around Coulomb's law and depth of understanding of electron configurations and the information it can convey. Begin new learning with Coulomb's law. Note that this law refers to all bodies of opposite charge. In this instance, the attraction, or Coulombic force, between all protons in the nucleus and an individual electron is estimated based on the radius between those bodies. Share Coulomb's law with students. For simplicity, explain that quantities q_1 and q_2, when referring to forces within the atom, can be considered as follows: q_1 = total number of protons while q_2 = the electron we are observing (always a value of 1 in this instance). The product of q_1 and q_2 is divided by the square of the radius between them. To check for understanding, ask students the following concerning the diagram used at the beginning of class. Answers in parenthesis) 1. For this diagram, what is the value of q_1? (7) 2. What do electrons A and B have in common? (they are both in the inner s orbital, they are very close to the nucleus)



Chapter 3 Simulation: Electron Configuration

Chapter 3 Practice Questions:
3.87–3.101

- For electron A, what values would we use for Coulomb's law? ($F_c = (7)(1)/r^2$ where r is a small value)
- For electron B, what values would we use for Coulomb's law? ($F_c = (7)(1)/r^2$ where r is a slightly larger value)
- Consider electrons A and B. Which would experience greater Coulombic force and why? (A because the top values of Coulomb's law are the same for A and B, but the denominator is smaller, so the F_c is larger)
- Compare electrons A, B, C, and D. Which electron would experience the least Coulombic force and why? (Electron D would experience lower Coulombic force because the distance r is the greatest. The larger the denominator, the smaller the value of F_c)
- Compare electrons A, B, C, and D. Which electron would take the most energy to remove? Why? (Electron A would take the greatest energy to remove because it is held in place with the greatest Coulombic force).
- Which electron(s) would you say are in the valence shell of this atom? (those in the 2s and 2p orbitals)
- Which electrons are core electrons? (those in the 1s orbital)

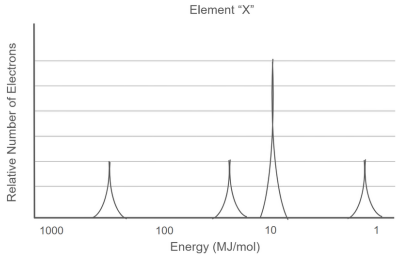
Once students have familiarity with comparing electrons based on Coulombic force, and after they have written justifications (see apply section below), they should practice writing electron configurations of neutral atoms from both diagrams like the one shown in the materials and from their periodic table. To see how the periodic table can be used as a guide, reference the YouTube video in the materials.

Once students master neutral atom electron configuration, expand the topic to ions. Returning to the first diagram, students can be asked to utilize an orbital filling diagram, like that found in the text. Ask students what is needed to fill the orbital. Remind them that the loss or gain of electrons allows for ions to form, holding ionic compounds together. In the case of the example, nitrogen can either lose its outer shell of 5 electrons or gain 3. It is more thermodynamically feasible to gain three, so the most common ion of nitrogen is N^{3-} . The electron configuration can show us that this has happened.

Apply New Knowledge

Have students practice translating between drawings of atoms, electron configurations, and orbital diagrams. Give them opportunities to justify explanations in writing using Coulomb's law. Chapter 3 "Electron Configuration" simulation can be used to support a visual representation of orbital filling as well as writing electron configurations. Provide opportunities for practice using problem sets such as those found in the materials.

Unit 1	Lesson Time	Suggested Skill	Learning Objective	Essential Knowledge
Topic 1.6 Photoelectron Spectroscopy	50 minutes	4.B	LO 1.6.A Explain the relationship between the photoelectron spectrum of an atom or ion and: i. The ground-state electron configuration of the species. ii. The interactions between the electrons and the nucleus.	EK 1.6.A.1

Materials	Instructional Activities and Assessments
<i>Chemistry: An Atoms-Focused Approach</i> Chapter 3, 3.11, pp. 127–131 AP Classroom Resource Exploring Atomic Structure Using Photoelectron Spectroscopy (PES) Data (Specifically parts 24 – 50) Make Connections Graphic  Photoelectron Spectroscopy Quiz: Practice Site Chapter 3 Practice Questions: 3.117 and 3.118	Activate Prior Knowledge For this lesson, draw on student's existing knowledge of electron configurations of neutral atoms and ions. For a warm up, show students a neutral electron configuration such as the one below and have them answer with shoulder partners. $1s^2 2s^2 2p^6 3s^2 3p^4$ Ask students the following: 1. Which element is shown in this electron configuration? (sulfur, it has 16 electrons and thus 16 protons) 2. How many valence electrons does this element have? How did you know? (6; the third energy level s and p orbital electrons add up to 6) 3. How many core electrons does this element have? How did you know? (10; electrons not in the valence shell are core electrons) 4. What is the term for the energy required to remove an electron from an atom? (ionization energy) 5. For which energy level or orbital in sulfur is the ionization energy lowest? Relate your answer to Coulomb's law. (The lowest ionization energy would be required for electrons in the third energy level s and p orbitals, specifically the 3p. These electrons are furthest from the nucleus, lowering the Coulombic force with which they are held.) 6. Complete this statement: The closer an electron is to the nucleus of an atom, the _____ the Coulombic force on the electron and the _____ the ionization energy required to remove the electron. (greater, greater) Make Connections to New Learning Photoelectron spectroscopy is one of the six types of spectroscopy tested on the AP® exam. It provides experimental evidence that the rules governing the writing of electron configurations is correct. Use the College Board resource linked in the materials to show students how photoelectron spectroscopy works (24 - 28 on the menu). Discuss how electrons are removed using X-ray or UV energy sources, and that the energy needed is equal to the ionization energy they have learned about. The more electrons removed at a certain energy, the higher the graphed peak on the diagram. Show students a PES diagram (see materials for an example). Describe each of the axes so that students understand that the height of the peak represents the number of electrons removed at the indicated energy. Connect the x-axis to

This is a preliminary sample; all content subject to change.

Coulombic force; the greater the value on the x-axis, the higher the ionization energy needed to remove that electron, the closer the electron was to the nucleus. Use the graphic and its relative placement of horizontal lines to help students see they can estimate the number of electrons based on the shortest peaks. If the leftmost peak represents the 2s and it is full, then the second horizontal line from the bottom must correspond to two electrons; $2s^2$. Energy levels can be identified as well by circling groups of peaks. Here the 1s can be circled, the 2s and 2p circled together, and the 3s circled. This can help students identify valence and core groupings. Have students practice writing electron configurations from PES graphs and identifying valence and core electrons as well.

Apply New Knowledge

Students can practice their skills with the text problems or quiz platform found in the materials.

Unit 1	Lesson Time	Suggested Skill	Learning Objective	Essential Knowledge
Topic 1.7 Periodic Trends	150 minutes	4.A	LO 1.7.A Explain the relationship between trends in atomic properties of elements and electronic structure and periodicity.	EK 1.7.A.1 EK 1.7.A.2 EK 1.7.A.3

Materials	Instructional Activities and Assessments
<i>Chemistry: An Atoms-Focused Approach</i> Chapter 2, 2.3, pp. 58–61 Chapter 3, 3.10–3.12, pp. 124–133 Chapter 4, 4.2, pp. 149–154 Simulation “Periodic Trends” Royal Society of Chemistry Interactive Periodic Table Chemistry: Argument using CER Chapter 3 and 4 Practice Questions: 3.107–3.122 and 4.13–4.18	Activate Prior Knowledge Before beginning new content, review Coulomb’s law and general periodic table information such as the location of metals and nonmetals, families/groups and periods, and family names (alkali metals, alkaline earth, halogens, etc.) The linked interactive periodic table in the materials may help with refamiliarizing students with these topics. After refreshing the idea of Coulomb’s law, remind students of the main components of Coulombic attraction; particulate radius and quantity of protons. Have students complete these statements to ensure their understanding of these relationships. - As the number of protons increases in an atom, the Coulombic force experienced by its electrons _____ (increases). - The closer an electron is to the nucleus, the _____ (greater) the Coulombic force it experiences. Make Connections to New Learning Each periodic trend can and should be related back to Coulombic force. Ensure students can explain the following trends using some of the key vocabulary (underlined terms) below. - Atomic/Ionic Radius - <u>Radius</u> is a measurement of the distance between the nucleus of an atom and its outermost electron cloud. - Atomic size decreases across a <u>period</u> on the periodic table. The electrons are being added to the same <u>valence</u> energy level and thus the major factor of change across a period is the increasing number of protons in the nucleus. An increase of protons leads to an increase in the numerator of Coulombs law, meaning an increase in <u>Coulombic force</u>. As Coulombic force increases, electrons are drawn closer to they nucleus and the overall size of the atom is decreased. - Atomic size increases down a <u>group/family</u> on the periodic table. Electrons are added to higher energy levels going down the group and thus are located farther from the nucleus. While the number of protons also increases, the equation value of the radius is squared, so the change in radius is more impactful. As the denominator grows, the <u>Coulombic force</u> decreases and electrons in the <u>valence</u> shell are held with less force, leading to an increase in the atomic size. <u>Cations</u> are formed when an atom loses electrons. Cations are smaller in radius than the <u>neutral</u> atoms from which they are formed. When a cation is formed, the number of <u>protons</u> remains consistent, meaning the <u>Coulombic</u> force remains consistent. The electrons that remain after some are removed are held by

- the same quantity of force. Electrons are removed from valence shells first, so the loss of an entire shell means the remaining electrons are found in the core energy levels, close to the nucleus.
- Anions are formed by the gain of electrons to fill a valence shell. While, again, the number of protons remains the same, the number of electrons increases, so the Coulombic force is distributed over a greater number of electrons, many of which are in the valence shell, far from the nucleus. This leads to an increase in radius when compared to the neutral atom.
 - Ionization Energy - Ionization energy is the energy expended to remove an electron from an atom or ion.
 - Ionization energy increases from left to right across a period. Valence electrons are located in the same energy level, but held with increasing numbers of protons, so the Coulombic force increases across a period. An electron held with more Coulombic force requires more ionization energy to remove from its atom.
 - One general exception to the rule occurs midway through the p-orbital elements. There are three p-orbitals on each period, starting at period 2. P-orbitals are filled with one up-spin electron, then one down-spin electron until each orbital holds two electrons. An exception to the ionization energy trend is seen between nitrogen and oxygen as well as phosphorus to sulfur. This occurs because the fourth electron added to the p-orbital becomes the first paired electron. Paired electrons repel one another, leading to a slightly higher orbital radius and a lowering of Coulombic force. This is not observed in the d or f orbitals as the energy level is so high that small changes to atomic radius due to electron pair repulsion are negligible.
 - Ionization energy decreases down a family/group. As valence shells are found on higher energy levels farther from the nucleus, the Coulombic force is less. This leads to lower ionization energies needed to remove valence electrons.
 - As a general note, when multiple ionizations occur, the ionization energy increases each time, so the first ionization energy is lower than the second which is lower than the third. This occurs because the number of protons and thus the Coulombic force is not changing, but is acting on fewer electrons each time, so it takes more energy to remove each one.
 - Removal of core electrons will always take significantly more ionization energy than removal of valence electrons due to the change in the radius component of Coulomb's law; as radius increases, Coulombic force increases and ionization energy increases.
 - Electron affinity - Electron affinity is the measure of the energy change that occurs when an electron is added to an atom or ion.
 - Elements that tend to lose electrons to form stable ions, like group 1 metals, tend to have less negative electron affinities. This indicates that the addition of an electron is not as thermodynamically favorable; it takes a great deal of energy to "force" another electron into the valence shell due to the low Coulombic force exhibited by the nucleus.
 - Elements that tend to gain electrons to form anions, like the halogens, tend to have high negative values of electron affinity. As they gain electrons, they become more isoelectronic to their nearest noble gas neighbor, and this increase in stability drive the gain of electrons.
 - Electronegativity - Electronegativity is a measure of the tendency of an atom to attract an electron from a shared, or bonding pair of electrons.
 - Electronegativity is a direct correlation to Coulombic force. Atoms with a great number of protons and valence electrons in lower energy shells closer to the nucleus have strong Coulombic attraction for their own electrons, as well as those that are shared from other atoms in molecular bonds. The greatest electronegativity values are found in the smallest halogens.

This is a preliminary sample; all content subject to change.

- Note: Noble gases tend to not form bonds, although a few of the larger atoms will form a few, due to their lack of electronegativity. Their full valence shells lead to great stability.

Apply New Knowledge

Have students complete the “Periodic Trends” simulation for practice and to better understand the correlation between Coulombic force and periodic trends. Have students work in small groups to practice AP Classroom problems, especially those that ask students to discuss and justify periodic trends. Students should use Claim, Evidence, Reasoning (CER or the CLEVER system) responses when writing answers to free response questions involving justification.

Note that the AP® exam does not test the drawing of orbital notations. However, these drawings are a great tool for determining how the distribution of electrons might affect another property and may be used as part of the justification in a free-response question.

This is a preliminary sample; all content subject to change.

Unit 1	Lesson Time	Suggested Skill	Learning Objective	Essential Knowledge
Topic 1.8 Valence Electrons and Ionic Compounds	40 minutes	4.C	LO 1.8.A Explain the relationship between trends in the reactivity of elements and periodicity.	EK 1.8.A.1 EK 1.8.A.2 EK 1.8.A.3

Materials	Instructional Activities and Assessments
<i>Chemistry: An Atoms-Focused Approach</i> Chapter 2, 2.3, pp. 58–61 Chapter 3, 3.9, pp. 121–123 Chapter 4, 4.2, pp. 149–156 Animation, Chapter 4, Bonding Chapter 3 and 4 Practice Problems: 2.11, 4.12 (d,e,f), 4.19–4.23	Activate Prior Knowledge Remind students that Coulombic forces lead to attraction between electrons and the nucleus of atoms. Utilize the periodic table to remind students of the common monatomic ion charges of the representative elements (group 1 metals have a charge of 1+ while group 7 nonmetals have a charge of 1-, for example). Review transition metals and remind students that their charges may vary due to a number of potential oxidation states. Finally, remind students of the position of the metals and nonmetals on the periodic table. Make Connections to New Learning Ensure students know that chemical bonding can occur in two ways: - When electrons are lost by one atom, gained by another, and two ions are formed, an ionic bond is created. In this type of bond, positive cations are attracted to negative anions, governed by Coulombic forces. Smaller ions tend to have stronger attractions, and this can affect their chemical behavior. - Note that elements in the same family will bond in the same way. For example, all group one alkali metals will form a +1 cation and will ionically bond in a 1:1 ration with group 7 halogens, which take on a 1- charge. - When electrons are shared by atoms with similar electronegativity, covalent (molecular) bonds are formed. If the two atoms in the bonding pair have the same electronegativity, then the electrons are shared equally and a nonpolar bond is formed. If the electronegativity of the bonding pair of atoms is unequal, then a polar bond is formed and the bonded pair will experience a dipole. - Utilize the bonding animation as you explain types of bonding and the movements of electrons to form ionic and covalent compounds. Apply New Knowledge Have students practice identifying the types of bonds that should exist in compounds based on their identity (metal + nonmetal is typically ionic while nonmetal + nonmetal is typically covalent) and electronegativity (a difference in electronegativity between two atoms greater than 1.7 is a good indicator of ionic bonding). Students should also be able to describe the range of polarity found in covalent bonds. Have students practice in small groups, focusing on justifications of the nature of bonds and the relationship of periodicity to types of bonds that occur between atoms.