# Ionic and Covalent Bonding

## Table of Contents

<table>
<thead>
<tr>
<th>Subject</th>
<th>Page Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Scenes 1-9 Introduction and Review Topics</td>
<td>3</td>
</tr>
<tr>
<td>Scenes 10-29 The Octet Rule and Ionic Bonds</td>
<td>7</td>
</tr>
<tr>
<td>Scenes 30-34 Properties of Ionic Compounds and Review of Ionic Bonding</td>
<td>13</td>
</tr>
<tr>
<td>Scenes 35-38 Metallic Bonds, Properties of Metals and Review of Metallic Bonds</td>
<td>16</td>
</tr>
<tr>
<td>Scenes 39-47 Covalent Bonds</td>
<td>18</td>
</tr>
<tr>
<td>Scenes 48-55 Properties of Covalent Bonds</td>
<td>22</td>
</tr>
<tr>
<td>Scenes 56-57 Review of Covalent Bonds and Conclusion</td>
<td>25</td>
</tr>
<tr>
<td>Multiple Choice Exam</td>
<td>27</td>
</tr>
<tr>
<td>Exam Answer Key</td>
<td>36</td>
</tr>
</tbody>
</table>
Scenes 1-9

Introduction and Review

I. Introduction...............................................................................................................1

II. Review Topics.............................................................................................................2

A. Energy levels and orbitals..........................................................................................2
B. Orbital diagram.............................................................................................................3
   1. Building an orbital diagram......................................................................................4
   2. Practice......................................................................................................................5
C. Electron configurations.................................................................................................6
   1. and the periodic table...............................................................................................7
   2. Practice......................................................................................................................8
D. Valence shells and electrons.........................................................................................9
Scene 1
You certainly notice differences between a gold ring, table salt, and the rubber sole of your shoe. Among other things, the three items differ in size, shape, and function. Do you notice any similarities among the three items? One of the most fundamental similarities cannot be seen by the naked eye. Each of these three items is a conglomeration of atoms resulting from bonds formed among the atoms. Depending on the atoms involved, three different types of bonds can occur: ionic bonds, metallic bonds, and covalent bonds. In this program you will see which elements participate in the different bond types, how each type of bond is formed, and the properties that each type of bond confers upon the compounds formed.

Scene 2
A quick review of electron orbitals and electron configuration will be beneficial for an understanding of chemical bonding. Within atoms, electrons may be found in seven energy levels. In each level, there are orbitals where electrons can be found. The letters s, p, d, and f represent orbitals. However, for the purposes of this review, only the s and p orbitals shall be discussed.

Scene 3
Orbital diagrams are used to indicate the placement of electrons in an atom at ground state. At ground state, all the electrons within an atom are at their lowest possible energy level. Therefore, electrons are placed in the diagram in a sequential manner. As you can see from the diagram, the orbitals are arranged by increasing energy. You should note that each energy level contains only one s orbital, indicated by a single solid line, and that the energy levels above level one contain three p orbitals, indicated by three short lines.

Scene 4
There are a few rules you should know prior to placing electrons in the orbital diagram. First, arrows representing electrons are placed in the orbitals one at a time. Second, always begin with the 1s orbital because it has the lowest energy. Third, no orbital can have more than two electrons, and the two electrons are displayed in opposite directions. Fourth, when more than one orbital of the same type occurs in a specific level, distribute the electrons one per orbital before adding the second electron to the orbital. Orbitals that contain only one electron are called half-filled orbitals, and are the orbitals that are important in bond formation. Fifth, work your way up the diagram so that when all the orbitals in an energy level are filled, electrons are placed in the next higher level.
Scene 5

Here’s how to draw an orbital diagram for fluorine, which is the ninth element, and thus has nine electrons. Energy level one contains only one orbital, designated as 1s. Consequently, the first two electrons fill the 1s orbital before other orbitals are filled. Only one s orbital occurs in the next higher energy level as well. Therefore, the next two electrons fill the line for the 2s orbital. The remaining five electrons fill the next higher orbitals, 2p. These electrons fill each orbital one at a time, before a second electron is added to any orbital. This results in a diagram with two filled 2p orbitals and one half-filled p orbital.

Scene 6

The arrangement of electrons in an atom is known as its electron configuration. Electron configurations can be established by referring to the orbital diagram created for a particular atom. Electron configurations are written in order from lowest energy orbitals to higher energy orbitals. For example, the electron configuration for fluorine is 1s² 2s² 2p⁵. The numbers represent the energy level in which the electrons are present. As you can see for the fluorine atom, electrons fall into the first and second energy levels. The letters indicate the orbital in which the electrons can be found. In this case, electrons travel in the s and p orbitals. The superscripts indicate how many electrons are present in each type of orbital. The s orbital in energy level one contains two electrons, the s orbital in energy level two contains two electrons, and the three p orbitals in level two contain a total of five electrons.

Scene 7

You can also use the periodic table as a quick guide to determine electron configuration for the elements contained within the first three periods. Note that the periodic table breaks into four blocks: s,p,d, and f. The order in which electron configurations are written is similar to the order in which you write, left to right and top to bottom. Starting at the top left hand side of the periodic table, work your way across the table filling in the electrons in their perspective orbitals along the way. Add one electron for each element crossed. Once finished with one line or level, move to the next energy level, and fill in the electrons as you work your way across it.

Scene 8

Using this method, try writing the electron configuration for an atom located in the third energy level. Phosphorus is element fifteen so it contains fifteen electrons. It has two electrons in the 1s block, two electrons in the two s block, six electrons in the 2 p block, two electrons in the 3s block and three electrons in the 3p block. If you have been counting, you know that the electron configuration for phosphorus is 1s² 2s² 2p⁶ 3s² 3p³. Lengthier electron configurations can be abbreviated by substituting the electron configuration of the noble gas that lies prior to the atom you are describing with the noble gas’s symbol. For example, the electron configuration for neon is 1s², 2s², 2p⁶.
Therefore, this portion of the configuration can be replaced with the symbol for neon in the electron configuration of phosphorous \([\text{Ne}]3s^23p^3\).

**Scene 9**

Electrons can be found in varying energy levels, called shells. An atom’s outermost shell is called the valence shell and the electrons located in this shell are appropriately called valence electrons. For simplicity, only the electrons in the valence shell are shown when discussing bonding because they are the electrons that are either donated or shared to form bonds. Atoms tend to gain, lose, or share electrons in order to achieve a complete valence shell.
Scenes 10-29
The Octet Rule and ionic Bonds

Scene Number

III. The Octet Rule................................................................................................................. 10

IV. Ionic Bonds........................................................................................................................ 11

A. Lewis dot diagrams............................................................................................................. 12
   2. Rearrangement of bonding electrons............................................................................. 13
      a. An example................................................................................................................ 14
   3. Practice............................................................................................................................ 15
B. Types of ions....................................................................................................................... 16
   1. Monatomic cations......................................................................................................... 17
      a. How to name them..................................................................................................... 18
   2. Monatomic anions.......................................................................................................... 19
      a. How to name them..................................................................................................... 20
   3. Polyatomic ions.............................................................................................................. 21
C. Binary ionic compounds...................................................................................................... 22
   1. Neutral compounds....................................................................................................... 23
   2. Empirical and chemical formulas.................................................................................. 24
      a. Determining ratios..................................................................................................... 25
         (1) Using the crisscross method............................................................................... 26
         (2) Checking your answer......................................................................................... 27
D. Ion ratios in polyatomic compounds.................................................................................. 28
   1. Checking your answer.................................................................................................... 29
Scene 10

The tendency for many elements to gain, lose, or share electrons so that the valence shell is filled is known as the octet rule. It is called the octet rule because the filled valence shell of these atoms contains eight electrons, and an octet means “set of eight”. Hydrogen and helium are excluded from this rule because their valence shell is in energy level one, which requires only two electrons to fill the valence shell. Reference to the octet rule will be made often during discussion of the three types of bonds.

Scene 11

The first type of bond discussed in this program is called an ionic bond. In ionic bonds, electrons in the valence shell are transferred between two atoms, completing the valence shell of each atom involved. Take a look at how sodium, a silvery metal, and chloride, a poisonous gas ionically bond to form sodium chloride, which you probably know as common table salt. Presented here is a representation of the valence electrons for both sodium and chlorine atoms in what is called a Lewis Dot Diagram. In Lewis dot diagrams, dots representing valence electrons are drawn around the element’s symbol. Again, the reason only valence electrons are shown is because they are the electrons involved in bonding. Therefore, the diagrams indicate the “eagerness” of an element to find a compatible partner that will help the element achieve a full valence shell, or octet.

Scene 12

Remember that the electron configuration for sodium is [Ne]3s\(^1\) and that the configuration for chlorine is [Ne]3s\(^2\)3p\(^5\). Notice how the abbreviated form of the electron configuration allows you to concentrate on only the valence electrons. The Lewis dot diagram for sodium is drawn with one dot, which represents its single valence electron, next to the sodium symbol, Na. The diagram for chlorine is drawn with seven dots representing valence electrons around the chlorine symbol, Cl. Each dot is placed sequentially around each side of the symbol. The maximum number of electrons allowed per side is two.

Scene 13

Lewis dot diagrams are useful because they illustrate the rearrangement of bonding electrons during chemical reactions. In order for these atoms to achieve a full set of valence electrons or an octet, the sodium atom must lose one electron and the chlorine atom must gain one electron. The process of electron transfer creates ions, which are charged atoms or charged groups of atoms. You will learn more about ions shortly.
Scene 14

In this equation, gray dots represent the sodium atom’s electron and green dots represent chloride ion’s electrons. Since electrons carry a negative charge, the sodium ion with a missing electron has a charge of $1^+$ (plus) and the chloride ion with an additional electron has a charge of $1^-$ (minus). Positive charges are denoted by the number of electrons the atom lost in the reaction followed by a plus sign. Negative charges are denoted by the number of electrons the atom gained in the reaction followed by a minus sign. Oppositely charged ions, like oppositely charged magnet ends, have such a strong attraction for one another that they come together, forming a bond that makes a neutral compound, meaning that the compound has no charge.

Scene 15

Try drawing a few more Lewis dot diagrams and reactions before moving on to the next topic. What is the electron configuration of magnesium? If you said (Ne)3s$^2$ you are correct! Therefore, placing two dots around the atomic symbol, Mg, makes the Lewis dot diagram for magnesium. What is the electron configuration for oxygen? (on screen) The resulting Lewis dot diagram is the symbol for oxygen, O, surrounded by six valence electrons. Do you think magnesium and oxygen could combine to form a compound? The Lewis diagram indicates that oxygen has six valence electrons and magnesium has two valence electrons. Magnesium donates its two valence electrons to oxygen, resulting in Mg$^{2+}$, and oxygen accepts the two electrons, becoming O$^{2-}$, satisfying the octet rule for each element. In addition, the attraction between the resulting oppositely charged ions results in a bond, forming the neutral compound magnesium oxide.

Scene 16

To this point you have only been introduced to the term ion, which refers to an atom or a group of atoms that possesses a charge. However, you probably noticed that ions are either positively or negatively charged. Positively charged ions, such as the calcium ion Ca$^{2+}$, are called cations. Negatively charged ions, such as the chloride ion Cl$^{-}$, are called anions. Cations and anions can be further divided into groups depending upon how many atoms combine to form the ion. The ions seen thus far in the program are called monatomic ions because the ions are formed by only one atom.

Scene 17

The elements in the two columns on the left-hand side of the periodic table form monatomic cations. Mono means one and atomic refers to the atom. Therefore monatomic cations are composed of only one atom. Those elements in group one must lose one electron to achieve a full electron shell and therefore form monatomic cations with a $1^+$ charge. The elements in group two form monatomic cations with a $2^+$ charge because they must lose two electrons to achieve a full valence shell. Lithium and beryllium...
are exceptions to the octet rule when forming ions because the loss of electrons results in the electron configuration of helium. Remember, two electrons fill helium's valence shell. Some elements form more than one kind of cation. The transition metals, which are located in the center of the periodic table, do not follow the octet rule when forming cations because electrons can occupy orbitals not involved in bonding. For example, it is common to find iron, which is denoted as Fe, as having of charge of 2+ or 3+.

**Scene 18**

Cations follow a consistent naming pattern. They always have the same name as the element. For example Mg$^{2+}$ is a magnesium ion and K$^+$ is a potassium ion. When elements that can have more than one charge are named using the Stock system, the element’s name is given first and then the charge is specified in Roman numerals. For example, copper$^+$ is written as copper (1) and copper$^{2+}$ is written as copper (II). See the chart provided to learn more common ions and their names.

**Scene 19**

The elements on the right of the periodic table, the nonmetals, form monatomic anions most easily. The charge of these anions can be determined by their position in the periodic table. The anions of nonmetals take on the charge equal to the number of elements they lie away from the nearest noble gas. For example, the anion of chlorine takes on the charge 1- while the anion for phosphorous takes on the charge 3-. The reason for equating the number from the nearest noble gas to the charge is because the elements must gain that many electrons to achieve a complete valence shell, thus fulfilling the octet rule.

**Scene 20**

The rule for naming monatomic anions is to replace the element's suffix with the letters i.d.e. However, no rule tells you where to place the new suffix. Bromine, for example, becomes bromide and oxygen becomes oxide. This chart shows you a few more names of monatomic anions.

**Scene 21**

In contrast to monatomic ions, polyatomic ions consist of more than one atom. A carbonate ion (CO$_3^{2-}$) represents a polyatomic anion. Here one carbon atom and three oxygen atoms bond to form the carbonate ion. While the bonds between carbon and oxygen in the carbonate ion are covalent and will be discussed later, the four atoms act as a unit that has an overall charge of 2-.

**Scene 22**

The cations created by metals and the anions created by nonmetals combine because their opposite charges attract one another. The simplest combination of ions occurs between two elements that produce monatomic ions of opposite charge, such as the sodium and chloride ions you saw earlier. The compounds formed are called binary ionic compounds because they form from the union of two different elements. “Bi” means two and “ionic” refers to the fact that the compounds are composed of ions. Binary ionic compounds are named by simply writing the cation element followed...
by the anion name. Therefore the combination of calcium and fluorine would yield the binary ionic compound calcium fluoride. What compound results from the combination of chlorine and potassium? If you said potassium chloride you are correct.

**Scene 23**

When cations and anions bond, their union creates a neutral, noncharged compound. This occurs when the ratio of cations to anions has the same magnitude, but opposite charge. In essence, the charges cancel each other. The mixture of potassium and chlorine illustrates this. Since the charge on the potassium ion is 1+ and the charge on the chlorine ion is 1-, a one to one ratio of each element neutralizes the charges, thus producing a neutral compound. However, not all ionic compounds occur in a one to one ratio of the atoms involved, such as the compound calcium fluoride CaF₂. Unfortunately, the chemical name of binary ionic compounds does not give any indication of the ratio in which atoms occur in ionic compounds.

**Scene 24**

Ratios of ions in ionic compounds are expressed in empirical or chemical formulas. Empirical formulas indicate the elements in a compound in their simplest whole ratio. To determine how to write these formulas, let’s take a look at calcium fluoride. The chemical formula for calcium fluoride is CaF₂. The elemental symbols represent the ions in the compound and the subscript represents how many ions of that element are present. An element that lacks a subscript behind its symbol has only one ion in the compound. Presented on the screen is another ionic compound, aluminum oxide.

**Scene 25**

You can use your recently acquired knowledge to determine the ratios of ions in an ionic compound. Using calcium oxide as an example, a list of steps showing how to determine the correct ratios of ions in a compound will be displayed. First, write the symbols of the elements found in the compound, remembering to write the cation’s symbol followed by the anion’s symbol. The symbol for calcium is Ca and the symbol for oxygen is O. Second, write the charge of each ion above its symbol. Refer to the periodic table if needed. The calcium cation forms a 2+ charge and the oxide anion forms a 2- charge.

**Scene 26**

Continuing the process, the third step is to move the charge of the cation to the subscript position of the anion. Also, move the charge of the anion to the subscript of the cation. In this case, the two from calcium is written behind O and the two from oxide is written behind Ca. This method of determining ion ratios is called the criss-cross method because the charge of one ion is criss-crossed to the subscript position of the other ion, and vice versa. In step four, make sure the subscripts occur in the smallest whole number possible. From the crisscross method, you see that calci
um oxide has two ions of each element. However, the smallest whole ratio of two-to-two is one-to-one. Finally, if a one is present as a subscript, it can be dropped since a lack of a subscript means that only one ion is present. Consequently, the empirical formula calcium oxide is written as CaO.

Scene 27

One way to check your answer is to make sure that the final compound is neutral. One calcium ion has a charge of 2+ and one oxide ion has a charge of 2-. Two + and 2- cancel each other, yielding a neutral compound. Also, it is good practice to make sure the ions in the compound satisfy the octet rule by drawing the Lewis dot structures. The calcium in this example loses its outer two electrons to oxygen, resulting in two ions that both attain a full valence shell and an ionic bond between them.

Scene 28

The crisscross method also works for determining the formula of polyatomic compounds. Remember that these compounds are composed of ions that contain more than one atom. Calcium phosphate is an example of an ionic compound that has a polyatomic anion. If needed, consult your textbook to find charges of polyatomic ions. One phosphorus atom and four oxygen atoms bond covalently to form a phosphate anion, which has a charge of 3-. You will learn about covalent bonds shortly. As you know, the calcium cation has a charge of 2+. Using the crisscross rule, move the charge of calcium behind the phosphate ion and move the charge for the phosphate anion behind the calcium symbol. At this step, parentheses are added around the phosphate ion to indicate that more than one polyatomic ion occurs in the compound. Next, make sure the subscripts occur in the smallest whole number ratio possible. Finally, drop any subscript that is a one, if applicable.

Scene 29

Once again, to double check that Ca₃(PO₄)₂ is the correct empirical formula, check for a neutral charge. Three calcium ions, each with a charge of 2+ are present. This brings the total charge for all calcium ions to 6+. Two phosphate ions, each with a charge of 3– are also present, bringing the total charge for phosphate ions to 6–. Since 6+ and 6– cancel, you can be confident that the formula is correct!
Scenes 30-34
Properties of Ionic Compounds and Review of Ionic Bonding

Scene Number

V. Properties of Ionic Compounds..................................................................................30
   A. Crystal lattice........................................................................................................30
   B. Melting point........................................................................................................31
   C. Hard, yet brittle character....................................................................................32
   D. Electrical conductivity.........................................................................................33

VI. Review of Ionic Bonding.........................................................................................34
Scene 30

Once ionic bonds form, the ions become tightly packed together because the oppositely charged ions strongly attract each other. Attraction between oppositely charged ions and repulsion between ions of the same charge are strong forces. Consequently, ions in an ionic compound align in an alternating pattern, cation, anion, cation, and so forth. This alternating arrangement of oppositely charged ions maximizes the attractions between opposite charges and minimizes the repulsions between like charges. This arrangement forms a three-dimensional structure called a crystal lattice. As a result of the crystal lattice structure, ionic compounds possess characteristic properties, which you will now explore.

Scene 31

Crystal lattice formation explains why ionic compounds are solids at room temperature and why high temperatures are required to melt them. At room temperature, ions within a compound are tightly packed. They cannot move about the lattice structure because the attractions are too strong. However, they do vibrate in place. When heated, the ions gain energy and begin to move within the lattice, bouncing off one another. At a certain temperature, the energy gained from heat is great enough to overcome the attractions between cations and anions. This temperature, which is different for each ionic compound, is called the melting point. Melting points for ionic compounds are very high because the bonds between cations and anions are very strong.

Scene 32

The crystal lattice also gives ionic compounds their properties of being hard, but brittle. Again, the bonds between the cations and anions are very strong, therefore, ionic compounds are difficult to crush and are not very flexible. However, if struck in a certain way, ionic compounds can fracture, shattering in a particular manner. When struck, ions in a crystal lattice structure shift, which can cause cations to align with cations and anions to align with anions. The repulsion between the ions with the same charge becomes so intense that the crystal shatters along the fissure where the ions shifted.

Scene 33

One property that the crystal lattice does not confer is electrical conductivity because electrical charges cannot move through the fixed ions in the lattice. Ionic compounds only conduct electricity when melted or dissolved in water because ions are then capable of moving. A simple experimental way to determine if a substance conducts electricity is to place the melted substance or dissolved solution into a container with one electrode hooked to a light bulb and another hooked to a power source. If an electrical current passes through the solution, then the light bulb shines. Here you see that melted and dissolved sodium chloride conduct an electrical current and solid sodium chloride does not.
Scene 34

A quick review of ionic bonds and compounds is in order before moving to the other types of bonds. Atoms often gain or lose electrons to achieve eight electrons in their valence shell. The exchange of electrons creates ions. Positively charged ions are due to the loss of electrons and negatively charged ions are due to the gain of electrons. Metals form cations and nonmetals form anions. The attraction between cations and anions is so great that the ions come together, forming an ionic bond. Ionic bonds are formed between a metal cation and nonmetal anion. Ionic bonds result in neutral ionic compounds that form a crystal lattice. The crystal lattice structure of ionic compounds confers many of the properties of ionic compounds, including their high melting points, and their hard yet brittle character.
**Scenes 35-38**  
Metallic Bonds, Properties of Metals and Review of Metallic Bonds

### Scene Number

<table>
<thead>
<tr>
<th>Scene</th>
<th>Description</th>
<th>Page</th>
</tr>
</thead>
<tbody>
<tr>
<td>VII.</td>
<td>Metallic Bonds</td>
<td>35</td>
</tr>
<tr>
<td></td>
<td>A. Electron sea mode</td>
<td>35</td>
</tr>
<tr>
<td>VIII.</td>
<td>Properties of Metals</td>
<td>36</td>
</tr>
<tr>
<td></td>
<td>A. Electric and heat conduction</td>
<td>36</td>
</tr>
<tr>
<td></td>
<td>B. Ductility and malleability</td>
<td>37</td>
</tr>
<tr>
<td>IX.</td>
<td>Review of Metallic Bonds</td>
<td>38</td>
</tr>
</tbody>
</table>
Scene 35

The second type of bond discussed in this program is the metallic bond. Unlike ionic bonds, which occur between metal and nonmetal atoms, metallic bonds occur strictly between metal atoms. The simplest model of metallic bonding is called the electron sea model. This model suggests that mobile valence electrons surround closely packed metal cations, forming a structure like marbles packed in a jar with BB’s. Instead of electrons being anchored to cations, represented here by the marbles, the electrons drift freely among cations throughout the structure, just as these BB’s are free to move around each marble.

Scene 36

The electron sea model also explains properties of metals, such as electric and heat conductivity, ductility, and malleability, which you will learn about soon. Electric current occurs due to the free movement of electrons. Metals make good electric conductors because the electrons constantly move throughout the metal. For the same reason, metals are also good conductors of heat. The movement of electrons allows the rapid passage of heat throughout the metal, because the electrons transfer thermal energy.

Scene 37

Ductility and malleability are two properties referring to the ability to shape metals. Ductile objects can be easily pulled into wires. Since copper can easily be drawn into wires and it conducts electricity well, copper wires are extensively used in computers, air conditioners, appliances, and other objects that use electricity. In contrast, malleable objects can be pounded and bent into shape. This property differs from the ionic crystal lattices that, as you remember, break along fissures. The reason why metals do not break like ionic crystals is because the electrons essentially act as ball bearings allowing the cations to slide past one another while reducing the repulsive force between the cations. The malleable properties of metals can be seen in gold medallions and other treasures.

Scene 38

A quick review of metallic bonds reminds you that only metal atoms are involved in metallic bonds. The electron sea model is used to describe the way in which metals bond. Electrons move freely among the various cations that are closely packed in a metal. This free movement of the electrons gives metals their characteristic properties, such as electrical and heat conductivity, malleability, and ductility.
Scenes 39-47
Covalent Bonds

Scene Number

X Covalent Bonds............................39

A. Molecular formula............................39
B. Lewis structures...............................40
   1. and valence electrons..................41
   2. Nonbonding valence electrons........42
   3. Multiple bonds............................43
      a. Double bonds.........................43
      b. Triple bonds............................44
   4. Exceptions to the octet rule.........45
      a. Fewer electrons than an octet...45
      b. More electrons than an octet.....46
   5. Resonance structures....................47
Scene 39

The third type of bond to be discussed is called a covalent bond. Unlike ionic bonds where atoms transfer electrons before forming bonds, electrons are shared between atoms that are covalently bonded. Two or more atoms linked together by covalent bonds are called molecules. The number and types of atoms that constitute a molecule are described by a molecular formula. For example, methane has a molecular formula of CH₄. This indicates that one carbon atom and four hydrogen atoms compose the methane molecule.

Scene 40

To show how atoms are covalently bonded to one another, chemists often use Lewis structures. Lewis structures resemble the Lewis dot diagrams used to show ionic bonding. However, dashes replace dots in order to show that electrons are being shared. Each dash represents two shared electrons in a bond, and counts as two valence electrons in each atom’s valence shell. Using methane as an example, first determine the number of valence electrons for each atom involved in the molecule, and draw each atom’s Lewis dot diagram. Remember that a hydrogen atom has one electron in its valence shell and that a carbon atom has four electrons in its valence shell. Also remember that a hydrogen atom requires one additional electron to fill its valence shell and a carbon atom requires four additional electrons to fill its valence shell.

Scene 41

To arrange atoms correctly within a molecule, make sure that each atom satisfies the octet rule by achieving a complete valence shell when bonded to other atoms in the molecule. Which example do you think shows the atoms arranged correctly in the methane molecule? The molecule on the right shows the correct arrangement of atoms. Remembering that a dash represents two electrons, you can see that each of the atoms has a complete valence shell. Two electrons are shared between each hydrogen and carbon atom, yielding the two electrons required to fill each hydrogen atoms’ valence shells and the eight electrons required to fill the oxygen’s valence shell. See how the molecule on the left would only satisfy the valence requirement for three hydrogen atoms in the molecule? As a result, the central hydrogen exceeds the correct number of valence electrons, making this an incorrect Lewis structure. In addition, the carbon atom is two electrons short of a full valence shell.

Scene 42

Not all electrons in the valence shell are required to bond to form molecules. Water has the molecular formula of H₂O. Drawing the six valence electrons around oxygen and the one valence electron around hydrogen, we can see how the hydrogen atoms bond to the oxygen atom. At this point the requirements to satisfy the valence shells for each atom are complete. However, two pairs
point the requirements to satisfy the valence shells for each atom are complete. However, two pairs of electrons on the oxygen atom are not shared with other atoms and are referred to as unshared electrons, a nonbonding pair, or a lone pair. Nonetheless, the valence shell for each atom is filled.

Scene 43

The examples of covalent bonding to this point have only shown a single bond between each atom in a molecule. In a single bond, two atoms share an electron pair. Many molecules have atoms that share more than one electron pair. Some atoms form double bonds, such as carbon and oxygen in carbon dioxide. The Lewis dot diagram shows the four valence electrons in the carbon atom and the six valence electrons in the oxygen atoms. To satisfy the octet rule for each atom, the carbon atom must form a double bond with each oxygen atom. Remember that each dash represents two electrons, therefore the four dashes represent eight electrons around the carbon atom. The oxygen atoms satisfy the octet rule with four shared electrons in the double bonds and four unshared electrons per atom.

Scene 44

Triple bonds also occur between atoms. Acetylene, a fuel burned in welding torches, has the molecular formula C₂H₂. How do these atoms combine to form a molecule that fills the valence shell for each atom? Placing the valence electrons for each atom around its atomic symbol, you can see that one hydrogen atom bonds with one carbon atom to complete the hydrogens’ valence shells. At this point, both carbons have five electrons in their valence shell, but they need eight to satisfy the octet rule. What bonding arrangement leads to eight electrons in each carbon’s valence shell? The answer is the sharing of the remaining electrons in each carbon atom by the formation of three bonds, which is called a triple bond.

Scene 45

As with most rules, there are exceptions. Of course there are some exceptions to the octet rule. Sometimes molecules violate the octet rule either by having more than an octet or less than an octet. Molecules containing the element boron often have fewer than eight electrons in boron’s valence shell. One example is boron trifluoride. As you can see in this Lewis structure, the boron atom has only six valence electrons – two electrons shy of a full valence shell. The three fluorine atoms, however, have complete valence shells. Molecules that do not fill their valence shell tend to be very reactive because they are “eager” to fill their valence shell.

Scene 46

Atoms such as sulfur and phosphorous sometimes form bonds that result in more than an octet of electrons in their valence shell. To demonstrate this, look at a sulfur hexafluoride molecule, SF₆. Each fluorine atom has seven electrons in the valence shell and sulfur has six electrons in its valence shell. Therefore, forty-eight electrons need to be accommodated within the molecule.
Bonding the sulfur atom to all the fluorine atoms satisfies the octet rule for the fluorine atoms. However, this also leads to the sulfur atom possessing twelve electrons. The reason these atoms can form bonds while possessing more than eight electrons is because the extra four electrons fill empty 3d orbitals.

**Scene 47**

Sometimes more than one Lewis structure illustrates the correct relationship between atoms. Ozone, which has the molecular formula of \( \text{O}_3 \), satisfies the octet rule when an oxygen atom forms one single bond with one oxygen atom and a double with the another oxygen atom. However, it is not completely accurate to depict ozone with this Lewis structure because the double bond could have formed between the other oxygen just as easily. These alternative structures are what chemists call resonance structures. They are drawn with a double-headed arrow between the structures indicating that the molecule’s true structure is not one or the other, rather it is an average or superposition of all possible resonance structures. In fact, experiments have shown that the bond lengths are shorter than the length of a single bond yet longer than the bond length for a double bond. The experimental data enforce the idea that the true structure is an average between the resonance structures for the molecule.
Scenes 48-55
Properties of Covalent Bonds

Scene Number

XI Properties of Covalent Bonds

A. Electronegativity
   1. and polar bonds
   2. and nonpolar bonds
   3. Predicting polarity
   4. Polar bonds in nonpolar molecules

B. Melting and boiling points
   1. of network solids
Scene 48
Molecules show diverse properties including molecular length and shape, states of matter, polarity, and melting and boiling points, all of which are due to the sharing of electrons in a covalent bond. Molecules may be simple and contain as few as two atoms, such as the oxygen molecule, or they may be complex and contain millions of atoms, such as the double helical DNA molecule that determines genetic inheritance. In contrast to ionic compounds, which you may remember are solid at room temperature, molecules may be found as gases, liquids, or solids, depending on the molecule. At room temperature oxygen is a gas, water is a liquid, and common table sugar is a solid. To learn more about how bonding affects these properties, consult your textbook or the CyberEd program Molecular Shape and Intermolecular Forces.

Scene 49
Electrons are not always shared equally between atoms in a molecule because some atoms are more electron “greedy” than others - that is, different atoms have different electronegativities. Electronegativity is the tendency for an atom to attract the shared electrons of a covalent bond more strongly to itself. This chart lists many elements and their electronegativities. Atoms with higher electronegativity values are more electron “greedy” than those atoms with smaller values.

Scene 50
When a large difference occurs between the electronegativities of two atoms forming a covalent bond, the bond is said to be polar. In polar bonds, the shared electron pair is strongly pulled to the more electronegative atom. For instance, polar bonds occur between the oxygen atom and the hydrogen atoms in a water molecule. The oxygen atom is more electronegative than the hydrogen atoms as the values in the chart indicate. Therefore, the oxygen atom tends to pull the electrons in each bond slightly towards itself. Consequently, the oxygen atom gains a partial negative charge and the hydrogen atoms gain a partial positive charge. The charges are noted with the Greek symbol delta. The atom with the slight negative charge is noted as delta minus and the atoms with the positive charge are the noted as delta plus.

Scene 51
In contrast, when the electronegativities of two atoms are similar or equal, nonpolar bonds form. In this type of bond, both atoms pull relatively equally on the electrons so the bond shows no polarity. Although fluorine is the most electronegative atom, the bond between two fluorine atoms is nonpolar because the difference in electronegativity between the atoms is zero.

Scene 52
This chart will help you predict
whether a bond is polar, nonpolar, or ionic. If the electronegativity difference between the two atoms is 0.4 or less, than the bond is a nonpolar covalent bond. Methane forms nonpolar bonds. If the difference is between 0.5 and 1.9 inclusive, the bond is a polar covalent bond, as is illustrated by hydrogen fluoride. If the difference is 2 or greater, the electrons are transferred and ionic bonds, which you learned about at the beginning of the program, form. Sodium chloride is an example of this. Take a moment to review these molecules, noting the difference in electronegativity values and the type of bond formed.

Scene 53

Do you think that molecules containing polar bonds are always polar molecules? Just because a molecule contains polar bonds does not necessarily mean the molecule is polar. Sometimes bond polarities cancel, rendering the molecule nonpolar overall. Carbon dioxide illustrates this nicely. Carbon dioxide contains two polar bonds, yet the molecule is nonpolar. The oxygen atoms are significantly more electronegative than the carbon atom, therefore, they pull the shared electrons more firmly towards themselves. However, the pulls occur in opposite directions, and therefore, essentially cancel each other. Polar molecules, such as this water molecule, have positive and negative charges at opposite ends.

Scene 54

Another property that varies due to the atoms bonded in a molecule are melting and boiling points. In comparison to ionic compounds, the melting and boiling points of most molecules are low. This is because the attractions between covalent molecules tend to be weaker than the attractions between ionic compounds. Consequently, not as much energy in the form of heat needs to be applied to break the bonds between molecules.

Scene 55

A few covalently bonded molecules occur as network solids, which require extremely high temperatures to melt. For example, diamond remains solid even when the temperature reaches 1064 degrees Celsius, a temperature high enough to melt this gold ring. Network solids are molecules in which each atom is covalently bonded to many other atoms, forming a sturdy structure. In a diamond, each carbon is bonded to four other carbons. In order to melt network solids, the covalent bonds within the network must be broken, which requires a great amount of energy in the form of heat. In contrast, melting nonnetwork solids requires heat energy to overcome attractions only between the molecules forming the solid. Therefore, nonnetwork solids melt at lower temperatures than network solids.
Scenes 56-57
Review of Covalent Bonds and Conclusion

Scene Number

XII. Review of Covalent Bonds..............................................................56

XIV. Conclusion....................................................................................57
Scene 56

This concludes the section on covalent bonding. The sharing of electrons in covalent bonds forms molecules. Covalent bonds allow each atom involved in the bond to fill their valence shell, thus satisfying the octet rule. The electronegativity value of atoms involved in covalent bonds determines the bond's polarity, but not the overall polarity of the molecule. Remember that bond polarities can cancel each other if the opposite charges of the polar bonds are in different directions. In general, molecules show diversity in many properties including size, shape, and melting points.

Scene 57

This concludes the program on ionic, metallic, and covalent bonds. You have learned that the ionic bonds formed between nonmetals and metals are very strong bonds. These strong ionic bonds in turn give ionic compounds their characteristic properties, such as crystal lattice structure and high melting points. Metallic bonds, which consist of only metal atoms, give metals their characteristic properties, including malleability and electric conductivity. Finally, covalent bonds form between non-metal atoms, creating molecules. Although they have weaker bonds than ionic compounds, molecules formed by covalent bonds show an incredible variation in properties, including size, shape, and physical properties.
Bonding I: Ionic and Covalent Bonding

Exam

1. Valence electrons are _____.
   A. found in the outermost energy level
   B. involved in bonding
   C. at ground state
   D. all of the above

2. The octet rule states that _____.
   A. atoms gain, lose, or share electrons to achieve a full valence shell
   B. compounds always consist of eight electrons
   C. molecules always consist of eight atoms
   D. all atoms require eight electrons to form bonds

3. Identify the correct order of events for ionic bonding.
   A. anion and cation bond, then transfer electrons
   B. atoms transfer electrons, then bond
   C. ions transfer electrons, then bond
   D. none of the above

4. Ionic compounds are composed of _____.
   A. a metal anion and a nonmetal cation
   B. two metal anions
   C. a metal cation and a nonmetal anion
   D. two nonmetal cations

5. Which of the following is an ionic compound?
   A. MgCl₂
   B. C₆H₁₂O₆
   C. Diamond
   D. CO₂
6. Which of the following is the correct Lewis dot diagram for nitrogen?

A.  

B.  

C.  

D.  

7. The correct electron configuration for nitrogen is _____.

A.  \([\text{He}]\; 2s^22p^3\)

B.  \([\text{Ne}]\; 2s^22p^3\)

C.  \([\text{N}]\; 2s^22p^3\)

D.  \([\text{C}]\; 2p^3\)

8. Which is the correct orbital diagram for boron?

A.  

B.  

C.  

D.  

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9. While taking into account the octet rule and the following Lewis dot diagrams, which of the following hypothetical elements would bond with this one oxygen atom?

A. 
B. 
C. 
D. 

10. Identify the correct Lewis dot diagram for silicon.

A. 
B. 
C. 
D. 

11. Ions are formed by ______?

A. the disintegration of one or more electrons
B. the sharing of one or more electrons between two atoms
C. the transfer of one or more electrons from one atom to another
D. the disintegration of one or more protons

12. Indicate the charge on the bromide ion in the equation resulting in the ionic compound potassium bromide. \( K^+ + Br \rightarrow KBr. \)

A. \( Br^+ \)
B. \( Br^- \)
C. \( Br^{2+} \)
D. \( Br^{2-} \)

13. Ionic compounds are ______.

A. always neutral
B. brittle
C. always composed of an equal proportion of cations to anions
D. both a and b
14. $\text{PO}_4^{3-}$ is a _____.

A. monatomic cation  
B. monatomic anion  
C. polyatomic anion  
D. polyatomic cation

15. How many fluorine atoms bond to calcium to form calcium fluoride?

A. one  
B. two  
C. three  
D. four

16. The binary ionic compound with the chemical formula $\text{AgCl}$ is called _____.

A. silver chlorine  
B. silver II chlorine  
C. silver chloride  
D. silver chloride I

17. Roman numerals in a cation's name indicates _____.

A. the number of valence electrons in the atom  
B. the number of electrons an atom gains when bonding to other atoms  
C. that the cation can form more than one cation  
D. both b and c

18. The carbonate ion ($\text{CO}_3^{2-}$) has an overall charge of 2-, so what is the empirical formula for calcium carbonate?

A. $\text{Ca}_2\text{CO}_3$  
B. $\text{CaCO}_3$  
C. $\text{Ca}_2(\text{CO}_3)_2$  
D. $\text{CaCO}$
19. The hypothetical ions $Z^{3+}$ and $Q^{2-}$ combine to form an ionic compound. In what proportions do these ions combine to produce a neutral compound?

A. $Z_6Q_6$
B. $Z_2Q_3$
C. $ZQ$
D. $Z_3Q_2$

20. The crystal lattice made by ionic compounds confers the following properties: _____.

A. high melting points
B. electrical conductivity
C. brittleness
D. A and C

21. The reason ionic compounds are brittle is because _____.

A. the bonds between the anions and cations are weak and therefore broken easily
B. a strong repulsive force is created when cations suddenly align
C. ionic compounds are actually very hard and not brittle
D. none of the above

22. In which of the following set-ups would an electric current be completed?

A. (ionic solid)
B. (NaCl in solution)
C. (molten NaCl)
D. both B and C

23. The electron sea model for metals suggests that _____.

A. valence electrons drift freely around the metal cations
B. valence electrons are anchored to specific metal cations
C. electrons idle around the metal cations
D. none of the above

24. Metals conduct both heat and electricity because _____.

A. atoms move freely within the metals, carrying heat and an electric current throughout a system
B. free moving electrons can carry both heat and electric current throughout a system
C. cations rotate around the stationary electrons
D. both A and C
25. The two properties of metals that pertain to their ability to be drawn into wires or pounded into shape are ______.

A. ductility and mallardability
B. ductility and malleability
C. conductivity and specific gravity
D. density and volatility

26. Covalent bonds form when ______.

A. two atoms transfer electrons
B. electrons are rotated around the atoms in the molecule
C. electrons are shared between two atoms
D. electrons exchange places between two atoms

27. A molecular formula indicates ______.

A. the type and number of atoms in a molecule
B. the smallest ratio of the atoms in a molecule
C. only the type of atoms involved in a molecule
D. the order in which atoms are arranged in a molecule

28. The number of electrons shared between two atoms forming a single covalent bond is ______.

A. one
B. two
C. three
D. four

29. Which of the following is the correct Lewis structure for ammonia (NH₃)?

A. 
B. 
C. 
D. 

A.  
B.  
C.  
D.  

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30. The placement of the double bond goes between which two atoms in this formaldehyde molecule?

A. Between the carbon atom and a hydrogen atom  
B. Between the carbon atom and the oxygen atom  
C. Between the two hydrogen atoms  
D. Between the oxygen atom and a hydrogen atom

31. Which of the following molecules does not obey the octet rule?

A. C₆H₁₂O₆  
B. CO₂  
C. SF₄  
D. All molecules obey the octet rule

32. These two Lewis structures for carbon dioxide satisfy the octet rule, and are therefore resonance structures.

A. True  
B. False
33. Which is the most electronegative element?

A. nitrogen  
B. fluorine  
C. sodium  
D. copper

34. Which of the following is a polar molecule?

A. water  
B. methane  
C. sodium chloride  
D. carbon dioxide

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<th>Element</th>
<th>Electronegativity</th>
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<td>Oxygen</td>
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<tr>
<td>Chlorine</td>
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<td>Carbon</td>
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<tr>
<td>Hydrogen</td>
<td>2.1</td>
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<tr>
<td>Sodium</td>
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</tbody>
</table>

A. Water  
B. Methane  
C. Sodium Chloride  
D. Carbon Dioxide

35. In a water molecule, the electrons within the molecule are attracted more towards the hydrogen atoms.

A. True  
B. False

36. All molecules that contain polar bonds are polar.

A. True  
B. False
37. Which of the following molecules possesses polar bonds, but is not a polar molecule?

A. methane  
B. sodium chloride  
C. carbon dioxide  
D. water

38. Network solids have very high melting points because _____.

A. the attraction between cations and anions is very strong  
B. the atoms within the solids are covalently bonded to four additional atoms  
C. submerging requires a great deal of heat energy  
D. many atoms are required to form network solids

39. Ionic compounds are named by writing the cation name, followed by the anion name.

A. True  
B. False
# Bonding I: Ionic and Covalent Bonding

## Exam Answer Key

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