

Chapter 6 Chemical Bonding

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Chapter 6 Section 1 Introduction to Chemical Bonding

Lesson Starter

- Imagine getting onto a crowded elevator. As people squeeze into the confined space, they come in contact with each other. Many people will experience a sense of being too close together.
- When atoms get close enough, their outer electrons repel each other. At the same time, however, each atom's outer electrons are strongly attracted to the nuclei of the surrounding atoms.
- The degree to which these outer electrons are attracted to other atoms determines the kind of chemical bonding that occurs between the atoms.

Chapter 6 Section 1 Introduction to Chemical Bonding

Objectives

- **Define** *chemical bond*.
- **Explain** why most atoms form chemical bonds.
- **Describe** ionic and covalent bonding.
- **Explain** why most chemical bonding is neither purely ionic nor purely covalent.
- **Classify** bonding type according to electronegativity differences.

Chapter 6 Visual Concepts

Chemical Bond

Ionic Bonding

Chapter 6 Visual Concepts

Covalent Bonding

Comparing Polar and Nonpolar Covalent Bonds

Chapter 6 Visual Concepts

Using Electronegativity Difference to Classify Bonding

Chapter 6

Section 1 Introduction to Chemical Bonding

Chemical Bonding, *continued*

Sample Problem A

Use electronegativity values listed in **Figure 20** from the previous chapter in your book, on page 161, and **Figure 2** in your book, on page 176, to classify bonding between sulfur, S, and the following elements: hydrogen, H; cesium, Cs; and chlorine, Cl. In each pair, which atom will be more negative?

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Objectives

- **Define** *molecule* and *molecular formula*.
- **Explain** the relationships among potential energy, distance between approaching atoms, bond length, and bond energy.
- **State** the octet rule.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Objectives, *continued*

- **List** the six basic steps used in writing Lewis structures.
- **Explain** how to determine Lewis structures for molecules containing single bonds, multiple bonds, or both.
- **Explain** why scientists use resonance structures to represent some molecules.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Molecular Compounds

- A **molecule** is a neutral group of atoms that are held together by covalent bonds.
- A chemical compound whose simplest units are molecules is called a **molecular compound**.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Molecular Compounds

- The composition of a compound is given by its chemical formula.
- A **chemical formula** indicates the relative numbers of atoms of each kind in a chemical compound by using atomic symbols and numerical subscripts.
- A **molecular formula** shows the types and numbers of atoms combined in a single molecule of a molecular compound.

Chapter 6

Visual Concepts

Chemical Formula

Structure of a Water Molecule

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Formation of a Covalent Bond

- Most atoms have lower potential energy when they are bonded to other atoms than they have as they are independent particles.
- The figure below shows potential energy changes during the formation of a hydrogen-hydrogen bond.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Formation of a Covalent Bond

- The electron of one atom and proton of the other atom *attract* one another.
- The two nuclei and two electrons *repel* each other.
- These two forces cancel out to form a covalent bond at a length where the potential energy is at a minimum.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

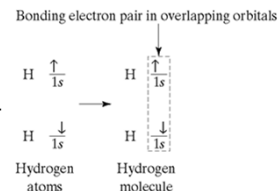
Characteristics of the Covalent Bond

- The distance between two bonded atoms at their minimum potential energy (the average distance between two bonded atoms) is the *bond length*.
- In forming a covalent bond, the hydrogen atoms release energy. The same amount of energy must be added to separate the bonded atoms.
- **Bond energy** is the energy required to break a chemical bond and form neutral isolated atoms.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Characteristics of the Covalent Bond

- When two atoms form a covalent bond, their shared electrons form overlapping orbitals.

- This achieves a noble-gas configuration.
- The bonding of two hydrogen atoms allows each atom to have the stable electron configuration of helium, $1s^2$.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

The Octet Rule

- Noble gas atoms are unreactive because their electron configurations are especially stable.
 - This stability results from the fact that the noble-gas atoms' outer *s* and *p* orbitals are completely filled by a total of eight electrons.
- Other atoms can fill their outermost *s* and *p* orbitals by sharing electrons through covalent bonding.
- Such bond formation follows the *octet rule*: Chemical compounds tend to form so that each atom, by gaining, losing, or sharing electrons, has an octet of electrons in its highest energy level.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

The Octet Rule, *continued* Exceptions to the Octet Rule

- Exceptions to the octet rule include those for atoms that cannot fit eight electrons, and for those that can fit more than eight electrons, into their outermost orbital.
 - Hydrogen forms bonds in which it is surrounded by only two electrons.
 - Boron has just three valence electrons, so it tends to form bonds in which it is surrounded by six electrons.
 - Main-group elements in Periods 3 and up can form bonds with *expanded valence*, involving *more* than eight electrons.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Electron-Dot Notation

- To keep track of valence electrons, it is helpful to use electron-dot notation.
- **Electron-dot notation** is an electron-configuration notation in which only the valence electrons of an atom of a particular element are shown, indicated by dots placed around the element's symbol. The inner-shell electrons are not shown.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Electron-Dot Notation, *continued*


Sample Problem B

- Write the electron-dot notation for hydrogen.
- Write the electron-dot notation for nitrogen.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Lewis Structures



- Electron-dot notation can also be used to represent molecules.
- 
- The pair of dots between the two symbols represents the shared electron pair of the hydrogen-hydrogen covalent bond.
 - For a molecule of fluorine, F_2 , the electron-dot notations of two fluorine atoms are combined.



Chapter 6

Section 2 Covalent Bonding and Molecular Compounds



Lewis Structures

- The pair of dots between the two symbols represents the shared pair of a covalent bond.
- 
- In addition, each fluorine atom is surrounded by three pairs of electrons that are not shared in bonds.
- 
- An *unshared pair*, also called a *lone pair*, is a pair of electrons that is not involved in bonding and that belongs exclusively to one atom.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Lewis Structures

- The pair of dots representing a shared pair of electrons in a covalent bond is often replaced by a long dash.
- example:  
- A **structural formula** indicates the kind, number, and arrangement, and bonds but not the unshared pairs of the atoms in a molecule.
- example: $F-F$ $H-Cl$

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Lewis Structures

- The Lewis structures and the structural formulas for many molecules can be drawn if one knows the composition of the molecule and which atoms are bonded to each other.
- A single covalent bond, or **single bond**, is a covalent bond in which one pair of electrons is shared between two atoms.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Lewis Structures, *continued*

Sample Problem C

Draw the Lewis structure of iodomethane, CH_3I .

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Multiple Covalent Bonds

- A double covalent bond, or simply a *double bond*, is a covalent bond in which two pairs of electrons are shared between two atoms.
 - Double bonds are often found in molecules containing carbon, nitrogen, and oxygen.
- A double bond is shown either by two side-by-side pairs of dots or by two parallel dashes.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Multiple Covalent Bonds

- A triple covalent bond, or simply a *triple bond*, is a covalent bond in which three pairs of electrons are shared between two atoms.
 - example 1—diatomic nitrogen:
 - example 2—ethyne, C_2H_2 :

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Multiple Covalent Bonds

- Double and triple bonds are referred to as **multiple bonds**, or multiple covalent bonds.
 - In general, double bonds have greater bond energies and are shorter than single bonds.
 - Triple bonds are even stronger and shorter than double bonds.
- When writing Lewis structures for molecules that contain carbon, nitrogen, or oxygen, remember that multiple bonds between pairs of these atoms are possible.

Chapter 6

Section 2 Covalent Bonding and Molecular Compounds

Multiple Covalent Bonds, *continued*

Sample Problem D

Draw the Lewis structure for methanal, CH_2O , which is also known as formaldehyde.

Chapter 6

Section 3 Ionic Bonding and Ionic Compounds

Objectives

- **Compare** a chemical formula for a molecular compounds with one for an ionic compound.
- **Discuss** the arrangements of ions in crystals.
- **Define** *lattice energy* and explain its significance.
- **List** and compare the distinctive properties of ionic and molecular compounds.
- **Write** the Lewis structure for a polyatomic ion given the identity of the atoms combined and other appropriate information.

Chapter 6

Section 3 Ionic Bonding and Ionic Compounds

Ionic Compounds

- Most of the rocks and minerals that make up Earth's crust consist of positive and negative ions held together by ionic bonding.
 - example: table salt, NaCl, consists of sodium and chloride ions combined in a one-to-one ratio— Na^+Cl^- —so that each positive charge is balanced by a negative charge.
- An **ionic compound** is composed of positive and negative ions that are combined so that the numbers of positive and negative charges are equal.

Chapter 6

Section 3 Ionic Bonding and Ionic Compounds

Ionic Compounds

- Most ionic compounds exist as crystalline solids.
- A crystal of any ionic compound is a three-dimensional network of positive and negative ions mutually attracted to each other.
- In contrast to a molecular compound, an ionic compound is not composed of independent, neutral units that can be isolated.

Chapter 6

Section 3 Ionic Bonding and Ionic Compounds

Ionic Compounds, *continued*

- The chemical formula of an ionic compound represents not molecules, but the simplest ratio of the compound's ions.
- A **formula unit** is the simplest collection of atoms from which an ionic compound's formula can be established.

Chapter 6

Section 3 Ionic Bonding and Ionic Compounds

Formation of Ionic Compounds

- The sodium atom has two valence electrons and the chlorine atom has seven valence electrons.
- Atoms of sodium and other alkali metals easily lose one electron to form cations.
- Atoms of chlorine and other halogens easily gain one electron to form anions.

Chapter 6

Section 3 Ionic Bonding and Ionic Compounds

Formation of Ionic Compounds, *continued*

- In an ionic crystal, ions minimize their potential energy by combining in an orderly arrangement known as a *crystal lattice*.
 - Attractive forces exist between oppositely charged ions within the lattice.
 - Repulsive forces exist between like-charged ions within the lattice.
- The combined attractive and repulsive forces within a crystal lattice determine:
 - the distances between ions.
 - the pattern of the ions' arrangement in the crystal

Chapter 6

Section 3 Ionic Bonding and Ionic Compounds

A Comparison of Ionic and Molecular Compounds

- The force that holds ions together in an ionic compound is a very strong electrostatic attraction.
- In contrast, the forces of attraction *between* molecules of a covalent compound are much weaker.
- This difference in the strength of attraction between the basic units of molecular and ionic compounds gives rise to different properties between the two types of compounds.

Chapter 6

Section 3 Ionic Bonding and Ionic Compounds

A Comparison of Ionic and Molecular Compounds, *continued*

- Molecular compounds have relatively weak forces between individual molecules.
 - They melt at low temperatures.
- The strong attraction between ions in an ionic compound gives ionic compounds some characteristic properties, listed below.
 - very high melting points
 - hard but brittle
 - not electrical conductors in the solid state, because the ions cannot move

Chapter 6

Section 3 Ionic Bonding and Ionic Compounds

Polyatomic Ions

- Certain atoms bond covalently with each other to form a group of atoms that has both molecular and ionic characteristics.
- A charged group of covalently bonded atoms is known as a **polyatomic ion**.
- Like other ions, polyatomic ions have a charge that results from either a shortage or excess of electrons.

Chapter 6

Section 3 Ionic Bonding and Ionic Compounds

Polyatomic Ions

- An example of a polyatomic ion is the ammonium ion: NH_4^+ . It is sometimes written as NH_4^+ to show that the group of atoms as a *whole* has a charge of 1+.
- Draw the Lewis structure for ammonium:
 - The seven protons in the nitrogen atom plus the four protons in the four hydrogen atoms give the ammonium ion a total positive charge of 1+.

Chapter 6

Section 4 Metallic Bonding

Objectives

- **Describe** the electron-sea model of metallic bonding, and explain why metals are good electrical conductors.
- **Explain** why metal surfaces are shiny.
- **Explain** why metals are malleable and ductile but ionic-crystalline compounds are not.

Chapter 6

Section 4 Metallic Bonding

Metallic Bonding

- Chemical bonding is different in metals than it is in ionic, molecular, or covalent-network compounds.
- The unique characteristics of metallic bonding gives metals their characteristic properties, listed below.
 - electrical conductivity
 - thermal conductivity
 - malleability
 - ductility
 - shiny appearance

Chapter 6

Section 4 Metallic Bonding

Metallic Bonding, *continued*

- **Malleability** is the ability of a substance to be hammered or beaten into thin sheets.
- **Ductility** is the ability of a substance to be drawn, pulled, or extruded through a small opening to produce a wire.

Chapter 6 Section 4 Metallic Bonding

The Metallic-Bond Model

- In a metal, the vacant orbitals in the atoms' outer energy levels overlap.
- This overlapping of orbitals allows the outer electrons of the atoms to roam freely throughout the entire metal.
- The electrons are *delocalized*, which means that they do not belong to any one atom but move freely about the metal's network of empty atomic orbitals.
- These mobile electrons form a *sea of electrons* around the metal atoms, which are packed together in a crystal lattice.

Chapter 6 Section 4 Metallic Bonding

The Metallic-Bond Model, *continued*

- The chemical bonding that results from the attraction between metal atoms and the surrounding sea of electrons is called **metallic bonding**.

Chapter 6 Visual Concepts

Metallic Bonding

Properties of Metals: Surface Appearance

Chapter 6 Visual Concepts

Properties of Metals: Malleability and Ductility

Properties of Metals: Electrical and Thermal Conductivity

Chapter 6 Section 5 Molecular Geometry

Objectives

- **Explain** VSEPR theory.

Chapter 6 Section 5 Molecular Geometry

Objectives, *continued*

- **Describe** dipole-dipole forces, hydrogen bonding, induced dipoles, and London dispersion forces and their effects on properties such as boiling and melting points.

Chapter 6 Section 5 Molecular Geometry

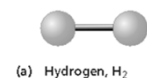
Molecular Geometry

- The properties of molecules depend not only on the bonding of atoms but also on *molecular geometry*: the three-dimensional arrangement of a molecule's atoms.
- The polarity of each bond, along with the geometry of the molecule, determines *molecular polarity*, or the uneven distribution of molecular shape.
- Molecular polarity strongly influences the forces that act *between* molecules in liquids and solids.
- A chemical formula, by itself, reveals little information about a molecule's geometry.

Chapter 6 Section 5 Molecular Geometry

VSEPR Theory

- As shown at right, diatomic molecules, like those of (a) hydrogen, H_2 , and (b) hydrogen chloride, HCl , can only be linear because they consist of only two atoms.



(a) Hydrogen, H_2



(b) Hydrogen chloride, HCl

- To predict the geometries of more-complicated molecules, one must consider the locations of all electron pairs surrounding the bonding atoms. This is the basis of VSEPR theory.

Chapter 6 Section 5 Molecular Geometry

VSEPR Theory

- The abbreviation VSEPR (say it "VES-pur") stands for "valence-shell electron-pair repulsion."
- VSEPR theory** states that repulsion between the sets of valence-level electrons surrounding an atom causes these sets to be oriented as far apart as possible.
 - example: BeF_2
 - The central beryllium atom is surrounded by only the two electron pairs it shares with the fluorine atoms.
 - According to VSEPR, the shared pairs will be as far away from each other as possible, so the bonds to fluorine will be 180° apart from each other.
 - The molecule will therefore be linear:

Chapter 6 Section 5 Molecular Geometry

VSEPR and Molecular Geometry

	Molecular shape	Atoms bonded to central atom	Lone pairs of electrons	Formula example	Lewis structure
Trigonal-pyramidal		3	1	NH_3	
Bent		2	2	H_2O	
Trigonal-bipyramidal		5	0	PCl_5	
Octahedral		6	0	SF_6	

Chapter 6 Section 5 Molecular Geometry

Hybridization

- Take the simple example of methane, CH_4 . The carbon atom has four valence electrons, two in the $2s$ orbital and two in $2p$ orbitals.
- How does carbon form four equivalent, tetrahedrally arranged, covalent bonds?
 - Recall that s and p orbitals have different shapes. To achieve four equivalent bonds, carbon's $2s$ and three $2p$ orbitals *hybridize* to form four new, identical orbitals called sp^3 orbitals.

Chapter 6 Section 5 Molecular Geometry

Hybridization, *continued*

- The four ($s + p + p + p$) hybrid orbitals in the sp^3 -hybridized methane molecule are equivalent: they all have the same energy, which is greater than that of the $2s$ orbital but less than that of the $2p$ orbitals.
- Hybrid orbitals** are orbitals of equal energy produced by the combination of two or more orbitals on the same atom.
- Hybridization explains the bonding and geometry of many molecules.

Chapter 6 Section 5 Molecular Geometry

Intermolecular Forces

- The forces of attraction between molecules are known as *intermolecular forces*.
 - The boiling point of a liquid is a good measure of the intermolecular forces between its molecules: the higher the boiling point, the stronger the forces between the molecules.
 - Intermolecular forces vary in strength but are generally weaker than bonds between atoms within molecules, ions in ionic compounds, or metal atoms in solid metals.
- Boiling points for ionic compounds and metals tend to be much higher than those for molecular substances: forces between molecules are weaker than those between metal atoms or ions.

Chapter 6 Section 5 Molecular Geometry

Intermolecular Forces, *continued*

- The strongest intermolecular forces exist between polar molecules.
- Because of their uneven charge distribution, polar molecules have *dipoles*. A **dipole** is created by equal but opposite charges that are separated by a short distance.
- The direction of a dipole is from the dipole's positive pole to its negative pole.

Chapter 6 Section 5 Molecular Geometry

Intermolecular Forces, *continued*

- A dipole is represented by an arrow with its head pointing toward the negative pole and a crossed tail at the positive pole. The dipole created by a hydrogen chloride molecule is indicated as follows:



Chapter 6 Section 5 Molecular Geometry

Intermolecular Forces, *continued*

- The negative region in one polar molecule attracts the positive region in adjacent molecules. So the molecules all attract each other from opposite sides.
- Such forces of attraction between polar molecules are known as *dipole-dipole forces*.
- Dipole-dipole forces act at short range, only between nearby molecules.
- Dipole-dipole forces explain, for example the difference between the boiling points of iodine chloride, I-Cl (97°C), and bromine, Br-Br (59°C).

Chapter 6 Section 5 Molecular Geometry

Intermolecular Forces, *continued*

- The small size of the hydrogen atom allows the atom to come very close to an unshared pair of electrons in an adjacent molecule.
- The intermolecular force in which a hydrogen atom that is bonded to a highly electronegative atom is attracted to an unshared pair of electrons of an electronegative atom in a nearby molecule is known as **hydrogen bonding**.

Chapter 6 Section 5 Molecular Geometry

Intermolecular Forces

- Hydrogen bonds are usually represented by dotted lines connecting the hydrogen-bonded hydrogen to the unshared electron pair of the electronegative atom to which it is attracted.
- An excellent example of hydrogen bonding is that which occurs between water molecules. The strong hydrogen bonding between water molecules accounts for many of water's characteristic properties.

Chapter 6

Section 5 Molecular Geometry

Intermolecular Forces, *continued* **London Dispersion Forces**

- Even noble gas atoms and nonpolar molecules can experience weak intermolecular attraction.
- In any atom or molecule—polar or nonpolar—the electrons are in continuous motion.
- As a result, at any instant the electron distribution may be uneven. A momentary uneven charge can create a positive pole at one end of an atom or molecule and a negative pole at the other.

Chapter 6

Section 5 Molecular Geometry

Intermolecular Forces, *continued* **London Dispersion Forces, *continued***

- This temporary dipole can then induce a dipole in an adjacent atom or molecule. The two are held together for an instant by the weak attraction between temporary dipoles.
- The intermolecular attractions resulting from the constant motion of electrons and the creation of instantaneous dipoles are called **London dispersion forces**.
- Fritz London first proposed their existence in 1930.