





## Chapter Summary and Review

### Key Learning Outcomes




CHAPTER OBJECTIVES	ASSESSMENT
<b>Classify Bonds as Pure Covalent, Polar Covalent, or Ionic</b> (5.2  )	<ul style="list-style-type: none"> <li>• Example 5.1  For Practice 5.1  Exercises 23 , 24 </li> </ul>
<b>Write Lewis Structures for Covalent Compounds</b> (5.3  )	<ul style="list-style-type: none"> <li>• Examples 5.2 , 5.3  For Practice 5.2 , 5.3  Exercises 27 , 28 , 29 , 30 </li> </ul>
<b>Write Lewis Structures for Polyatomic Ions</b> (5.3  )	<ul style="list-style-type: none"> <li>• Example 5.4  For Practice 5.4  Exercises 31 , 32 , 33 , 34 </li> </ul>
<b>Write Resonance Lewis Structures</b> (5.4  )	<ul style="list-style-type: none"> <li>• Example 5.5  For Practice 5.5  Exercises 35 , 36 </li> </ul>
<b>Assign Formal Charges to Assess Competing Resonance Structures</b> (5.4  )	<ul style="list-style-type: none"> <li>• Example 5.6  For Practice 5.6  For More Practice 5.6  Exercises 37 , 38 , 39 , 40 </li> </ul>
<b>Draw Resonance Structures and Assign Formal Charge for Organic Compounds</b> (5.4  )	<ul style="list-style-type: none"> <li>• Example 5.7  For Practice 5.7  Exercises 41 , 42 , 43 , 44 </li> </ul>
<b>Write Lewis Structures for Compounds Having Expanded Octets</b> (5.5  )	<ul style="list-style-type: none"> <li>• Example 5.8  For Practice 5.8  For More Practice 5.8  Exercises 47 , 48 , 49 , 50 </li> </ul>
<b>Use VSEPR Theory to Predict the Basic Shapes of Molecules</b> (5.7  )	<ul style="list-style-type: none"> <li>• Example 5.9  For Practice 5.9  Exercises 53 , 54 </li> </ul>
<b>Predict Molecular Geometries Using VSEPR Theory and the Effects of Lone Pairs</b> (5.8  )	<ul style="list-style-type: none"> <li>• Examples 5.10 , 5.11  For Practice 5.10 , 5.11  Exercises 55 , 56 , 57 , 58 </li> </ul>
<b>Predict the Shapes of Larger Molecules</b> (5.8  )	<ul style="list-style-type: none"> <li>• Example 5.12  For Practice 5.12  Exercises 63 , 64 , 67 , 68 </li> </ul>
<b>Use Molecular Shape to Determine Polarity of a Molecule</b> (5.10  )	<ul style="list-style-type: none"> <li>• Example 5.13  For Practice 5.13  Exercises 71 , 72 , 73 , 74 </li> </ul>

## Key Terms

### Section 5.2

polar covalent bond   
electronegativity   
dipole moment ( $\mu$ )   
percent ionic character 



### Section 5.4

resonance structure   
resonance hybrid   
formal charge 








### Section 5.5

free radical 









### Section 5.6

bond energy   
bond length 

### Section 5.7

valence shell electron pair repulsion (VSEPR) theory   
electron groups   
linear geometry   
trigonal planar geometry   
tetrahedral geometry   
trigonal bipyramidal geometry   
octahedral geometry 

### Section 5.8

electron geometry   
molecular geometry   
trigonal pyramidal geometry   
bent geometry   
seesaw geometry   
T-shaped geometry   
square pyramidal geometry   
square planar geometry 

## Key Concepts

### Morphine: A Molecular Impostor (5.1)

- The properties of molecules are directly related to their shapes.
- Many biological functions, such as drug action and the immune response, are determined by the shapes of molecules.

## Electronegativity and Bond Polarity (5.2)

- The shared electrons in a covalent bond are not always *equally* shared; when two dissimilar nonmetals form a covalent bond, the electron density is greater on the more electronegative element. The result is a polar bond, with one element carrying a partial positive charge and the other a partial negative charge.
- Electronegativity—the ability of an atom to attract electrons to itself in chemical bonding—increases as we move across a period to the right in the periodic table and decreases as we move down a column.
- Elements with very dissimilar electronegativities form ionic bonds; those with very similar electronegativities form nonpolar covalent bonds; and those with intermediate electronegativity differences form polar covalent bonds. The degree of polarity of a bond depends on the electronegativity difference between the bonding atoms.

## Writing Lewis Structures of Molecular Compounds and Polyatomic Ions (5.3)

- Follow the procedure in [Section 5.3](#) to draw Lewis structures for compounds.
- Electrons must be added (anions) or removed (cations) to account for the charge of a polyatomic ion.

## Resonance and Formal Charge (5.4)

- Some molecules are best represented not by a single Lewis structure, but by two or more resonance structures. The actual structure of these molecules is a resonance hybrid: a combination or average of the contributing structures.
- The formal charge of an atom in a Lewis structure is the charge the atom would have if all bonding electrons were shared equally between bonding atoms.
- In general, the best Lewis structures will have the fewest atoms with formal charge, and any negative formal charge will be on the most electronegative atom.

## Exceptions to the Octet Rule (5.5)

- Although the octet rule is normally used in drawing Lewis structures, some exceptions occur.
- These exceptions include odd-electron species, which necessarily have Lewis structures with only seven electrons around an atom. Such molecules, called free radicals, tend to be unstable and chemically reactive.
- Other exceptions to the octet rule include molecules with incomplete octets—usually just six electrons (especially important in compounds containing boron)—and molecules with expanded octets—usually 10 or 12 electrons (which can occur in compounds containing elements from the third row of the periodic table and below). Expanded octets never occur in second-period elements.

## Bond Energies and Bond Lengths (5.6)

- The bond energy of a chemical bond is the energy required to break 1 mole of the bond in the gas phase. Average bond energies for a number of different bonds are tabulated.
- Average bond lengths are also tabulated. In general, triple bonds are shorter and stronger than double bonds, which are in turn shorter and stronger than single bonds.

## VSEPR Theory: Predicting Molecular Shape (5.7–5.9)

- In VSEPR theory, molecular geometries are determined by the repulsions between electron groups on the central atom. An electron group can be a single bond, double bond, triple bond, lone pair, or even a single electron.
- The five basic molecular shapes are linear (two electron groups), trigonal planar (three electron groups), tetrahedral (four electron groups), trigonal bipyramidal (five electron groups), and octahedral (six electron groups).
- When lone pairs are present on the central atom, the *electron* geometry is still one of the five basic shapes, but one or more positions are occupied by lone pairs. The *molecular* geometry is therefore different from the electron geometry. Lone pairs are positioned so as to minimize repulsions with other lone pairs and with bonding pairs.
- To determine the geometry of a molecule, follow the procedure presented in [Section 5.9](#).

## Molecular Shape and Polarity (5.10)

- The polarity of a polyatomic molecule containing polar bonds depends on its geometry. If the dipole moments of the polar bonds are aligned in such a way that they cancel one another, the molecule will not be polar. If they are aligned in such a way as to sum together, the molecule will be polar.
- Highly symmetric molecules tend to be nonpolar, while asymmetric molecules containing polar bonds tend to be polar. The polarity of a molecule dramatically affects its properties.

## Key Equations and Relationships

**Dipole Moment ( $\mu$ ):** Separation of Two Particles of Equal but Opposite Charges of Magnitude  $q$  by a Distance  $r$  (5.2 □)

$$\mu = qr$$

**Percent Ionic Character** (5.2 □)

$$\text{percent ionic character} = \frac{\text{measured dipole moment of bond}}{\text{dipole moment if electron were completely transferred}} \times 100\%$$

**Formal Charge** (5.4 □)

$$\text{formal charge} = \text{number of valence electrons} - \left( \text{number of nonbonding electrons} + \frac{1}{2} \text{ number of shared electrons} \right)$$

Not for Distribution

*Not for Distribution*