CHAPTER 6 REVIEW
Chemical Bonding

SECTION 2

SHORT ANSWER  Answer the following questions in the space provided.

1. Use the concept of potential energy to describe how a covalent bond forms between two atoms.

   As the atoms involved in the formation of a covalent bond approach each other, the electron-proton attraction is stronger than the electron-electron and proton-proton repulsions. The atoms are drawn to each other and their potential energy decreases. Eventually, a distance is reached at which the repulsions between the like charges equals the attraction of the opposite charges. At this point, potential energy is at a minimum and a stable molecule forms.

2. Name two elements that form compounds that can be exceptions to the octet rule.

   Choose from hydrogen, boron, beryllium, phosphorus, sulfur, and xenon.

3. Explain why resonance structures are used instead of Lewis structures to correctly model certain molecules.

   Resonance structures show that one Lewis structure cannot correctly represent the location of electrons in a bond. Resonance structures show delocalized electrons, while Lewis structures depict electrons in a definite location.

4. Bond energy is related to bond length. Use the data in the tables below to arrange the bonds listed in order of increasing bond length, from shortest bond to longest.

   a.  
   
<table>
<thead>
<tr>
<th>Bond</th>
<th>Bond energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H—F</td>
<td>569</td>
</tr>
<tr>
<td>H—I</td>
<td>299</td>
</tr>
<tr>
<td>H—Cl</td>
<td>432</td>
</tr>
<tr>
<td>H—Br</td>
<td>366</td>
</tr>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>
   H—F, H—Cl, H—Br, H—I
SECTION 2 continued

b. | Bond | Bond energy (kJ/mol) |
---|---|---|
     | C—C | 346  |
     | C=C | 835  |
     | C≡C | 612  |
     | C≡C, C=C, C—C |

5. Draw Lewis structures to represent each of the following formulas:

a. \( \text{NH}_3 \)
   \[
   \begin{array}{c}
   \text{H} \\
   \text{N} \\
   \text{H} \\
   \text{H}
   \end{array}
   \]

b. \( \text{H}_2\text{O} \)
   \[
   \begin{array}{c}
   \text{H} \\
   \text{O} \\
   \text{H}
   \end{array}
   \]

c. \( \text{CH}_4 \)
   \[
   \begin{array}{c}
   \text{H} \\
   \text{H} \\
   \text{C} \\
   \text{H}
   \end{array}
   \]

d. \( \text{C}_2\text{H}_2 \)
   \[
   \begin{array}{c}
   \text{H} \\
   \text{C≡C} \\
   \text{H}
   \end{array}
   \]

e. \( \text{CH}_2\text{O} \)
   \[
   \begin{array}{c}
   \text{H} \\
   \text{H} \\
   \text{C}=\text{O}
   \end{array}
   \]
CHAPTER 6 REVIEW

Chemical Bonding

SECTION 3

SHORT ANSWER  Answer the following questions in the space provided.

1. a The notation for sodium chloride, NaCl, stands for one
   (a) formula unit.  (c) crystal.
   (b) molecule.   (d) atom.

2. d In a crystal of an ionic compound, each cation is surrounded by a number of
   (a) molecules.  (c) dipoles.
   (b) positive ions.  (d) negative ions.

3. b Compared with the neutral atoms involved in the formation of an ionic compound, the
   crystal lattice that results is
   (a) higher in potential energy.  (c) equal in potential energy.
   (b) lower in potential energy.  (d) unstable.

4. b The lattice energy of compound A is greater in magnitude than that of compound B. What
   can be concluded from this fact?
   (a) Compound A is not an ionic compound.
   (b) It will be more difficult to break the bonds in compound A than those in compound B.
   (c) Compound B has larger crystals than compound A.
   (d) Compound A has larger crystals than compound B.

5. b The forces of attraction between molecules in a molecular compound are generally
   (a) stronger than the attractive forces among formula units in ionic bonding.
   (b) weaker than the attractive forces among formula units in ionic bonding.
   (c) approximately equal to the attractive forces among formula units in ionic bonding.
   (d) equal to zero.

6. Describe the force that holds two ions together in an ionic bond.
   The force of attraction between unlike charges holds a negative ion and a positive
   ion together in an ionic bond.

7. What type of energy best represents the strength of an ionic bond?
   lattice energy
**SECTION 3 continued**

8. What types of bonds are present in an ionic compound that contains a polyatomic ion?

   The atoms in a polyatomic ion are held together with covalent bonds, but polyatomic ions combine with ions of opposite charge to form ionic compounds.

9. Arrange the ionic bonds in the table below in order of increasing strength from weakest to strongest.

<table>
<thead>
<tr>
<th>Ionic bond</th>
<th>Lattice energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCl</td>
<td>−787</td>
</tr>
<tr>
<td>CaO</td>
<td>−3384</td>
</tr>
<tr>
<td>KCl</td>
<td>−715</td>
</tr>
<tr>
<td>MgO</td>
<td>−3760</td>
</tr>
<tr>
<td>LiCl</td>
<td>−861</td>
</tr>
</tbody>
</table>

KCl, NaCl, LiCl, CaO, MgO

10. Draw Lewis structures for the following polyatomic ions:
   a. \( \text{NH}_4^+ \)

   ![Lewis structure for NH4+]

   b. \( \text{SO}_4^{2-} \)

   ![Lewis structure for SO42-]

11. Draw the two resonance structures for the nitrite anion, \( \text{NO}_2^- \).

   ![Two resonance structures for NO2-]
CHAPTER 6 REVIEW

Chemical Bonding

SECTION 4

SHORT ANSWER  Answer the following questions in the space provided.

1. **b** In metals, the valence electrons are considered to be
   (a) attached to particular positive ions.   (c) immobile.
   (b) shared by all surrounding atoms.    (d) involved in covalent bonds.

2. **a** The fact that metals are malleable and ionic crystals are brittle is best explained in terms of their
   (a) chemical bonds.   (c) enthalpies of vaporization.
   (b) London forces.   (d) polarity.

3. **d** As light strikes the surface of a metal, the electrons in the electron sea
   (a) allow the light to pass through.
   (b) become attached to particular positive ions.
   (c) fall to lower energy levels.
   (d) absorb and re-emit the light.

4. **d** Mobile electrons in the metallic bond are responsible for
   (a) luster.   (c) electrical conductivity.
   (b) thermal conductivity.   (d) All of the above.

5. **a** In general, the strength of the metallic bond moving from left to right on any row of the periodic table.
   (a) increases
   (b) decreases
   (c) remains the same
   (d) varies

6. **c** When a metal is drawn into a wire, the metallic bonds
   (a) break easily.
   (b) break with difficulty.
   (c) do not break.
   (d) become ionic bonds.

7. Use the concept of electron configurations to explain why the number of valence electrons in metals tends to be less than the number in most nonmetals.

   **Most metals have their outer electrons in s orbitals, while nonmetals have their outer electrons in p orbitals.**
8. How does the behavior of electrons in metals contribute to the metal’s ability to conduct electricity and heat?

The mobility of electrons in a network of metal atoms contributes to the metal’s ability to conduct electricity and heat.

9. What is the relationship between the enthalpy of vaporization of a metal and the strength of the bonds that hold the metal together?

The amount of energy required to vaporize a metal is a measure of the strength of the bonds that hold the metal together. The greater a metal’s enthalpy of vaporization, the stronger the metallic bond.

10. Draw two diagrams of a metallic bond. In the first diagram, draw a weak metallic bond; in the second, show a metallic bond that would be stronger. Be sure to include nuclear charge and number of electrons in your illustrations.

a.  

![weak bond](image1)

b.  

![strong bond](image2)

Note: In the strong bond, the charge on the nucleus and the number of electrons must be greater than in the weak bond.

11. Complete the following table:

<table>
<thead>
<tr>
<th></th>
<th>Metals</th>
<th>Ionic Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>Components</td>
<td>atoms</td>
<td>ions</td>
</tr>
<tr>
<td>Overall charge</td>
<td>neutral</td>
<td>neutral</td>
</tr>
<tr>
<td>Conductive in the solid state</td>
<td>yes</td>
<td>no</td>
</tr>
<tr>
<td>Melting point</td>
<td>low to high</td>
<td>high</td>
</tr>
<tr>
<td>Hardness</td>
<td>soft to hard</td>
<td>hard</td>
</tr>
<tr>
<td>Malleable</td>
<td>yes</td>
<td>no</td>
</tr>
<tr>
<td>Ductile</td>
<td>yes</td>
<td>no</td>
</tr>
</tbody>
</table>
SHORT ANSWER  Answer the following questions in the space provided.

1. Identify the major assumption of the VSEPR theory, which is used to predict the shape of atoms.
   
   **Pairs of valence electrons repel one another.**

2. In water, two hydrogen atoms are bonded to one oxygen atom. Why isn’t water a linear molecule?
   
   **The electron pairs that are not involved in bonding also take up space, creating a tetrahedron of electron pairs and making the water molecule angular or bent.**

3. What orbitals combine together to form \( sp^3 \) hybrid orbitals around a carbon atom?
   
   **the \( s \) orbital and all three \( p \) orbitals from the second energy level**

4. What two factors determine whether or not a molecule is polar?
   
   **electronegativity difference and molecular geometry or unshared electron pairs**

5. Arrange the following types of attractions in order of increasing strength, with 1 being the weakest and 4 the strongest.
   
   \[ 3 \] hydrogen bonding
   \[ 4 \] ionic
   \[ 2 \] dipole-dipole
   \[ 1 \] London dispersion

6. How are dipole-dipole attractions, London dispersion forces, and hydrogen bonding similar?
   
   **They are all forces of attraction between molecules. In all cases there is an attraction between the slightly negatively-charged portion of one molecule and the slightly positively charged portion of another molecule.**
7. Complete the following table:

<table>
<thead>
<tr>
<th><strong>Formula</strong></th>
<th><strong>Lewis structure</strong></th>
<th><strong>Geometry</strong></th>
<th><strong>Polar</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>H₂S</td>
<td><img src="image" alt="H₂S Lewis structure" /></td>
<td>bent</td>
<td>yes</td>
</tr>
<tr>
<td>CCl₄</td>
<td><img src="image" alt="CCl₄ Lewis structure" /></td>
<td>tetrahedral</td>
<td>no</td>
</tr>
<tr>
<td>BF₃</td>
<td><img src="image" alt="BF₃ Lewis structure" /></td>
<td>trigonal planar</td>
<td>no</td>
</tr>
<tr>
<td>H₂O</td>
<td><img src="image" alt="H₂O Lewis structure" /></td>
<td>bent</td>
<td>yes</td>
</tr>
<tr>
<td>PCl₅</td>
<td><img src="image" alt="PCl₅ Lewis structure" /></td>
<td>trigonal bipyramidal</td>
<td>no</td>
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<td>BeF₂</td>
<td><img src="image" alt="BeF₂ Lewis structure" /></td>
<td>linear</td>
<td>no</td>
</tr>
<tr>
<td>SF₆</td>
<td><img src="image" alt="SF₆ Lewis structure" /></td>
<td>octahedral</td>
<td>no</td>
</tr>
</tbody>
</table>