Two more classification schemes that might be used are: models that have the same elements bonding (O₂), and models that have different elements bonding (HCl); models with a carbon atom (CH₄), and models without a carbon atom (H₂O). (Students are likely to discover quickly the variety of compounds that can be formed using carbon atoms.) The rationale for this classification is that once again, clear distinctions can be made between such bonding arrangements.

### 2.1 CLASSIFYING COMPOUNDS

#### PRACTICE

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**Understanding Concepts**

1. (a) A metal element and a nonmetal element combine to form an ionic compound.
   (b) A nonmetal element and a nonmetal element combine to form a molecular compound.
2. Electrical conductivity can be used as a diagnostic test for an ionic compound. Ionic compounds (many of which dissolve readily in water) form solutions that conduct electricity.

**Applying Inquiry Skills**

3. **Experimental Design**
   - Solubility: Obtain a small amount of the unknown substance. Observe and record its state at the ambient temperature. Add a small quantity of the substance to about 10 mL of distilled water. Stir the mixture with a stirring rod and note whether the chemical dissolves. Many ionic compounds readily dissolve in water.
   - Conductivity: Obtain a small sample of distilled water in a beaker. Use a low-voltage conductivity apparatus to test the electrical conductivity of the sample. The apparatus should indicate a reading of zero. Test the electrical conductivity of the mixture from the above solubility procedure and record observations. Ionic compounds form solutions that conduct electricity.
4. - Compound A is ionic — its solution conducts electricity.
   - Compound B is molecular — it is a liquid at SATP, and its solution does not conduct electricity.
   - Compound C is molecular — it is a gas at SATP.
   - Compound D is ionic — its solution conducts electricity.
   - Compound E is molecular — its solution does not conduct electricity.

### 2.2 IONIC BONDING

#### PRACTICE

(Page 71)

**Understanding Concepts**

1. The properties of ionic compounds that suggest ionic bonds are strong are: they are solids at SATP, they have hard surfaces, and they have high melting and boiling points.
2. Metal elements and nonmetal elements form ionic bonds with each other.
3. Groups 1, 2, and 3 (13), tend to lose electrons to become positive ions. Groups 15, 16, and 17 tend to gain electrons to form negative ions.
4. The minimum number of different ions in the formula of an ionic compound is 2. This is because the smallest unit of an ionic compound that would still have the properties of the compound is a 1:1 ratio of the different ions, with the general formula MX.
5. (a) S²⁻  (f) K⁺
   (b) Ba²⁺  (g) P³⁻
   (c) Br⁻  (h) Rb⁺
   (d) Cl⁻  (i) Be²⁺
   (e) Ca²⁺
Applying Inquiry Skills

6. (a) **Experimental Design**

Conductivity of an Ionic Solid: Obtain a small sample of ionic solid such as a piece of chalk (calcium carbonate, CaCO₃(s)). Use a low-voltage conductivity apparatus to test the electrical conductivity of the sample. Record your observations.

Alternatively, fill a 100-mL beaker about one-third full with crystals of sodium chloride. Use a low-voltage conductivity apparatus to test the electrical conductivity of the sample. Record your observations.

(b) The student is to research the conductivity of molten (liquid) ionic compounds. The student will discover that in liquid form, ionic compounds conduct electricity.

(c) Ionic compounds are solids at room temperature. As solids they are nonconductors of electricity. However, as liquids they conduct electricity quite well. Some examples of molten ionic compounds are sodium chloride (melts at 801°C), and aluminum oxide, Al₂O₃ (melts at 2000°C).

(d) In the solid state, ionic compounds do not conduct electricity. In ionic solids, the ions are tightly held in the crystal structure, so they are not free to move and carry an electrical current. When an ionic compound is melted, the attractive forces are overcome and the crystal collapses. The ions are now free to move and carry an electrical current, as they are in solution.

**PRACTICE**

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**Making Connections**

7. The student is to use the Internet to research and report upon the importance of one of the ions that make up the human body, and comment on whether supplements of the ion are recommended. There are many ions from which to choose.

GO TO www.science.nelson.com, Chemistry 11, Teacher Centre.

**PRACTICE**

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**Understanding Concepts**

8. (a) The electron dot diagrams of metal ions differ from those of nonmetal ions in that the diagrams of metal ions have vacant valence orbitals, and no lone pairs, and show a positive charge outside the square bracket, while the diagrams of nonmetal ions have no vacant valence orbitals, have lone pairs, and show a negative charge outside the square bracket.

(b) Electron dot diagrams of metal ions are similar to those of nonmetal ions in that both diagrams show a total charge that is equal in value but opposite in sign, and both diagrams represent an ion that has full outer orbits of electrons — a configuration exactly the same as that of the nearest noble gas.

9. (a) lithium iodide \[ \text{Li}^+ \cdot \text{I}^- \rightarrow \text{[Li]}^+ [\text{I}]^- \]

   \[ \text{Li} + \text{I} \rightarrow \text{[Li]}^+ \text{[I]}^- \]

(b) barium chloride \[ \text{Ba}^+ \cdot 2\text{Cl}^- \rightarrow \text{[Ba]}^{2+} [\text{Cl}]^- \]

(c) potassium oxide \[ 2\text{K}^+ \cdot \text{O}^- \rightarrow \text{[K]}^+ [\text{O}]^2- [\text{K}]^- \]

(d) calcium fluoride \[ \text{Ca}^+ \cdot 2\text{F}^- \rightarrow \text{[Ca]}^{2+} [\text{F}]^- \]

10. (a) nitrogen \[ \text{N} \cdot \]

    (b) sulfur \[ \text{S} \cdot \]

    (c) argon \[ \text{Ar} \cdot \]

    (d) iodine \[ \text{I} \cdot \]

    (e) lithium \[ \text{Li} \cdot \]

    (f) cesium \[ \text{Cs} \cdot \]

    (g) calcium \[ \text{Ca} \cdot \]

    (h) sodium \[ \text{Na} \cdot \]
11. (a) \( \text{Ca} + \text{O} \rightarrow [\text{Ca}]^{2+} [\text{O}]^{2-} \)

The two elements will combine in a ratio of 1:1. The formula is CaO.

(b) \( 2 \text{Rb} + \text{O} \rightarrow [\text{Rb}]^{+} [\text{O}]^{2-} [\text{Rb}]^{+} \)

The two elements will combine in a ratio of 2:1. The formula is Rb₂O.

(c) \( \text{Sr} + \text{O} \rightarrow [\text{Sr}]^{2+} [\text{O}]^{2-} \)

The two elements will combine in a ratio of 1:1. The formula is SrO.

(d) \( 2 \text{Al} + 3 \text{O} \rightarrow [\text{O}]^{2-} [\text{Al}]^{3+} [\text{O}]^{2-} [\text{Al}]^{3+} [\text{O}]^{2-} \)

The two elements will combine in a ratio of 2:3. The formula is Al₂O₃.

12. \( \begin{array}{ccc}
\text{F} & : & \text{Cl} \\
: & : & : \\
\text{Br} & : & \text{I} \\
\end{array} \)

The electron dot diagrams for the five halogens are identical. Elements in the same chemical family have the same number of valence electrons, and will thus be represented by the same electron dot diagram.

13. (a) A magnesium atom has two valence electrons. By transferring the two electrons to chlorine atoms, the resulting magnesium ion will have a stable octet with the same electron configuration as neon.

A chlorine atom has seven valence electrons. By attracting an electron from a magnesium atom, the resulting chlorine ion will have a stable octet with the same electron configuration as argon. Two chlorine atoms will attract one electron each from a single magnesium atom.

\( \text{Mg} + 2 \text{Cl} \rightarrow [\text{Cl}]^- [\text{Mg}]^{2+} [\text{Cl}]^- \quad \cdot \text{Mg}^- + 2 \cdot [\text{Cl}]^- \rightarrow [\text{Mg}]^{2+} [\text{Cl}]^- \)

The formula is MgCl₂.

(b) A sodium atom has one valence electron. By transferring this electron to a sulfur atom, the resulting sodium ion will have a stable octet with the same electron configuration as neon.

A sulfur atom has six valence electrons. By attracting two electrons from sodium atoms, the resulting sulfur ion will have a stable octet with the same electron configuration as argon. One sulfur atom will attract one electron each, from two separate sodium atoms.

\( 2 \text{Na} + \text{S} \rightarrow [\text{Na}]^+ [\text{S}]^{2-} [\text{Na}]^+ \quad 2 \cdot \text{Na}^- + 2 \cdot [\text{S}]^- \rightarrow [\text{Na}]^+ [\text{S}]^{2-} [\text{Na}]^+ \)

The formula is Na₂S.

(c) An aluminum atom has three valence electrons. By transferring the three electrons to oxygen atoms, the resulting aluminum ion will have a stable octet with the same electron configuration as neon.

An oxygen atom has six valence electrons. By attracting two electrons from aluminum atoms, the resulting oxygen ion will have a stable octet, also with the same electron configuration as neon. Three oxygen atoms will attract two electrons each, from two separate aluminum atoms.

\( 2 \text{Al} + 3 \text{O} \rightarrow [\text{O}]^{2-} [\text{Al}]^{3+} [\text{O}]^{2-} [\text{Al}]^{3+} [\text{O}]^{2-} \quad 2 \cdot \text{Al}^- + 3 \cdot [\text{O}]^- \rightarrow [\text{O}]^{2-} [\text{Al}]^{3+} [\text{O}]^{2-} [\text{Al}]^{3+} [\text{O}]^{2-} \)

The formula is Al₂O₃.

(d) A barium atom has two valence electrons. By transferring the two electrons to chlorine atoms, the resulting barium ion will have a stable octet with the same electron configuration as xenon.

A chlorine atom has seven valence electrons. By attracting an electron from a barium atom, the resulting chlorine ion will have a stable octet with the same electron configuration as argon. Two chlorine atoms will attract one electron each from a single barium atom.

\( \text{Ba} + 2 \text{Cl} \rightarrow [\text{Cl}]^- [\text{Ba}]^{2+} [\text{Cl}]^- \quad \cdot \text{Ba}^- + 2 \cdot [\text{Cl}]^- \rightarrow [\text{Ba}]^{2+} [\text{Cl}]^- \)

The formula is BaCl₂.

(e) A calcium atom has two valence electrons. By transferring the two electrons to fluorine atoms, the resulting calcium ion will have a stable octet with the same electron configuration as xenon.

A fluorine atom has seven valence electrons. By attracting an electron from a calcium atom, the resulting fluorine ion will have a stable octet with the same electron configuration as neon. Two fluorine atoms will attract one electron each from a single calcium atom.

\( \text{Ca} + 2 \text{F} \rightarrow [\text{F}]^- [\text{Ca}]^{2+} [\text{F}]^- \quad \cdot \text{Ca}^- + 2 \cdot [\text{F}]^- \rightarrow [\text{F}]^- [\text{Ca}]^{2+} [\text{F}]^- \)

The formula is CaF₂.
(f) A sodium atom has one valence electron. By transferring this electron to an iodine atom, the resulting sodium ion will have a stable octet with the same electron configuration as neon.

An iodine atom has seven valence electrons. By attracting an electron from a sodium atom, the resulting iodine ion will have a stable octet with the same electron configuration as xenon.

\[
\text{Na} + \text{F} \rightarrow [\text{Na}]^+ [\text{F}]^-
\]

The formula is NaF.

(g) A potassium atom has one valence electron. By transferring this electron to a chlorine atom, the resulting potassium ion will have a stable octet with the same electron configuration as argon.

A chlorine atom has seven valence electrons. By attracting an electron from a potassium atom, the resulting chlorine ion will have a stable octet with the same electron configuration as argon.

\[
\text{K} + \text{Cl} \rightarrow [\text{K}]^+ [\text{Cl}]^-
\]

The formula is KCl.

14. (a) baking soda  
(b) table salt  
(c) limestone or chalk  
(d) slaked lime — used to make mortar and plaster

SECTION 2.1- 2.2 QUESTIONS

(Page 74)

Understanding Concepts

1. With respect to periodic trends, elements within a chemical family (group) tend to participate in similar chemical reactions, producing ionic compounds with the same general formula. With respect to electronegativity, Group 1 and 2 metals, which have low electronegativities, will readily react with the elements in Group 17, which have high electronegativities, to form ionic compounds. As most Group 1 and 2 metals and Group 17 elements are relatively abundant in nature, it makes sense that the ionic compounds that these elements form would also be abundant in nature.

2. To reach a stable state with a full outer orbit of electrons, oxygen must gain two electrons. To reach a stable state with a full outer orbit of electrons, lithium must lose one electron. Thus, for lithium and oxygen to combine, there must be two lithium atoms for every one oxygen atom, with each lithium atom donating one electron to a single oxygen atom — a ratio of two lithium atoms for every one oxygen atom.

3. Ion formation shows periodic trends. Elements within a chemical family (group) tend to produce similar ions, and to participate in similar chemical reactions, producing ionic compounds with the same general formula.

4. (a) calcium carbonate  
(b) calcium hydroxide  
(c) sodium chloride  
(d) sodium bicarbonate

5. (a) Cs + F → [Cs]^+ [F]^–  
(b) CsF  
(c) An ionic compound (an ionic halide).  
(d) Solid at SATP, with a hard surface, is brittle, has a high melting point, and forms a solution that conducts electricity.  
(e) The properties of the compound are due to the strong, simultaneous forces of attraction between the positive and negative ions, which hold the ions firmly in a rigid structure. The solid state, hardness, brittleness, and the high melting point result from the strong attractions, which occur in the crystal structure. And because the ionic bonds break down in water, the resulting ions are free to move in solution and conduct electricity.  
(f) Cs has a low first ionization energy, and F has a high electron affinity, so it makes sense that the reaction would be a vigorous one. With the alkali metals, reactivity increases moving down the group, due to increasing atomic radii and decreasing first ionization energies. With the halogens, reactivity increases moving up the group, due to decreasing atomic radii and increasing electron affinity.
2.3 COVALENT BONDING

PRACTICE

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Understanding Concepts

1. (a) $F_2(g)$ \[
\begin{align*}
\hat{F} + \hat{F} &\rightarrow \hat{F} \cdots \hat{F} \\
&\quad F \cdots F
\end{align*}
\]

(b) $H_2O(l)$ \[
\begin{align*}
H + O + H &\rightarrow O \cdots H \\
&\quad H \cdots O \cdots H
\end{align*}
\]

(c) $CH_4(g)$ \[
\begin{align*}
\hat{C} + H + H + H + H &\rightarrow C \cdots H \\
&\quad H \cdots C \cdots H \\
&\quad H \cdots H
\end{align*}
\]

(d) $PCl_3(s)$ \[
\begin{align*}
\hat{P} + \hat{Cl} + \hat{Cl} + \hat{Cl} &\rightarrow P \cdots Cl \\
&\quad Cl \cdots Cl \\
&\quad P \cdots Cl \\
&\quad Cl
\end{align*}
\]

(e) $H_2S(g)$ \[
\begin{align*}
H + \hat{S} + H &\rightarrow S \cdots H \\
&\quad H \cdots S \cdots H
\end{align*}
\]

(f) $SiO_2(s)$ \[
\begin{align*}
\hat{O} + \hat{Si} + \hat{O} &\rightarrow O = Si = O \\
&\quad O = Si = O
\end{align*}
\]

PRACTICE

(Page 79)

Understanding Concepts

2. (a) $H_2(g)$ \[
\begin{align*}
H + H &\rightarrow H \cdots H \\
&\quad H \cdots H
\end{align*}
\]

(b) $O_3(g)$ \[
\begin{align*}
\hat{O} + \hat{O} + \hat{O} &\rightarrow O \cdots O \cdots O \\
&\quad O \cdots O \cdots O
\end{align*}
\]

(c) $OF_2(g)$ \[
\begin{align*}
\hat{F} + \hat{O} + \hat{F} &\rightarrow F \cdots O \cdots F \\
&\quad F \cdots O \cdots F
\end{align*}
\]

(d) $NF_3(g)$ \[
\begin{align*}
\hat{N} + \hat{F} + \hat{F} + \hat{F} &\rightarrow N \cdots F \\
&\quad F \cdots F \\
&\quad F \cdots F
\end{align*}
\]

(e) $N_2H_2(g)$ \[
\begin{align*}
\hat{N} + H + \hat{N} + H &\rightarrow N \cdots N \cdots H \\
&\quad H \cdots N \cdots N \\
&\quad H \cdots H
\end{align*}
\]

(f) $P_2H_4(g)$ \[
\begin{align*}
\hat{P} + \hat{P} + \hat{H} + \hat{H} + \hat{H} + \hat{H} &\rightarrow P \cdots P \\
&\quad P \cdots P \\
&\quad P \cdots P \\
&\quad P \cdots P
\end{align*}
\]
3. (a) $\text{PO}_4^{3-} (\text{aq})$

\[
\begin{array}{c}
\text{O} \quad \text{P} \quad \text{O} \\
\text{O} \quad \text{P} \quad \text{O} \\
\text{O} \quad \text{P} \quad \text{O} \\
\end{array}
\]

(b) $\text{OH}^- (\text{aq})$

\[
\left[\text{O}^- \cdot \text{H} \rightarrow \left[\text{O} \cdot \text{H}^-\right]\right]
\]

(c) $\text{BrO}_3^- (\text{aq})$

\[
\left[\text{O} \cdot \text{Br} \cdot \text{O}^- \rightarrow \left[\text{O} \cdot \text{Br} \cdot \text{O}^-\right]\right]
\]

(d) $\text{ClO}_4^- (\text{aq})$

\[
\left[\text{O} \cdot \text{Cl} \cdot \text{O}^- \rightarrow \left[\text{O} \cdot \text{Cl} \cdot \text{O}^-\right]\right]
\]

4. Ozone, question 2(b), can be considered to have a coordinate covalent bond, with one of the oxygens starting as $[\text{O}]^{2-}$. Question 3 (a), (c), and (d) have coordinate covalent bonds; (b) would have a coordinate covalent bond if the ions involved are $[\text{O}]^{2-}$ and $[\text{H}]^{+}$.

**PRACTICE**

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**Understanding Concepts**

5. The term "bonding electrons" describes a pair of electrons that are shared by two atoms forming a covalent bond, whereas the term "lone pairs" describes a pair of electrons not involved in bonding.

6. (a) covalent
   (b) covalent
   (c) ionic
   (d) ionic

7. (a) Molecular elements and compounds: $\text{N}_2 (\text{g})$, $\text{O}_2 (\text{g})$, $\text{F}_2 (\text{g})$, $\text{C}_1 \text{H}_2 \text{O}_{11} (\text{sugar})$, $\text{C}_6 \text{H}_6 (\text{l})$ (liquid benzene), $\text{NH}_3 (\text{g})$ (ammonia).

   Ionic compounds: $\text{NaCl (s)}$, $\text{MgO (s)}$, $\text{Ca(OH)_2 (s)}$, $\text{CaF}_2 (\text{s})$, $\text{NaHCO}_3 (\text{s})$, $\text{CaCO}_3 (\text{s})$.

   (b) Molecular elements and compounds are formed by covalent bonds. Molecular compounds may be solids, liquids, or gases at SATP and tend to be soft or waxy. Covalent bonds between the atoms are strong. However, the intermolecular forces in molecular compounds are weaker in comparison — adding a relatively small amount of heat will cause a solid molecular compound to change state from a solid to a liquid, and then to a gas. Ionic compounds are formed by ionic bonds. Ionic compounds are solids at SATP and are hard and brittle. The properties of ionic compounds are due to the strong, simultaneous forces of attraction between the positive and negative ions, which hold the ions firmly in a rigid structure.

8. (a) The bonding capacity of nitrogen is three, and the bonding capacity of chlorine is one.
The number of covalent bonds (shared electron pairs) that an atom can form is known as its bonding capacity. Each atom of nitrogen shares three electron pairs, so it has a bonding capacity of three. Each atom of chlorine shares one electron pair, so it has a bonding capacity of one.

Coordinate covalent bonds are similar to covalent bonds in that once the bond is formed, there is no way to tell the difference from a covalent bond. Coordinate covalent bonds are different from covalent bonds in that instead of having the shared electrons come from both atoms, both of the shared electrons come from the same atom.

Two nitrogen atoms require three electrons each to form stable octets. The two atoms form a triple bond by sharing three electron pairs.

The strength of a bond between two atoms increases as the number of electron pairs in the bond increases. The triple bond in a nitrogen molecule is very strong, preventing the atoms from forming a bond with another atom.

### Bonding Capacities of Some Common Atoms

<table>
<thead>
<tr>
<th>Atom</th>
<th>Number of valence electrons</th>
<th>Number of bonding electrons</th>
<th>Bonding capacity</th>
</tr>
</thead>
<tbody>
<tr>
<td>carbon</td>
<td>4</td>
<td>4</td>
<td>4</td>
</tr>
<tr>
<td>nitrogen</td>
<td>5</td>
<td>3</td>
<td>3</td>
</tr>
<tr>
<td>oxygen</td>
<td>6</td>
<td>2</td>
<td>2</td>
</tr>
<tr>
<td>hydrogen</td>
<td>1</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>fluorine</td>
<td>7</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>chlorine</td>
<td>7</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>bromine</td>
<td>7</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>iodine</td>
<td>7</td>
<td>1</td>
<td>1</td>
</tr>
</tbody>
</table>

### Examples

(a) \( \text{HCl} \)  
(b) \( \text{NH}_3 \)  
(c) \( \text{H}_2\text{S} \)

HCl: \( \text{H}^+ + \text{Cl}^- \rightarrow \text{H} \equiv \text{Cl}^- \)

NH\(_3\): \( \text{N}^3+ \cdot \text{H}^+ \cdot \text{H}^+ \rightarrow \text{N} \equiv \text{H} \equiv \text{H} \)

H\(_2\)S: \( \text{S}^2+ + \cdot \text{H}^+ + \cdot \text{H}^+ \rightarrow \text{S} \equiv \text{H} \equiv \text{H} \)

(b) \( \text{Br}_2 \)  
(c) \( \text{H}_2\text{O}_2 \)  
(d) \( \text{C}_2\text{H}_4 \)

Br\(_2\): \( \text{Br} \equiv \text{Br} \)

H\(_2\)O\(_2\): \( \text{H} \equiv \text{O} \equiv \text{O} \equiv \text{H} \)

C\(_2\)H\(_4\): \( \text{H} \equiv \text{C} \equiv \text{C} \equiv \text{H} \)

### Formulas

- HCl: \( \text{H}^+ + \text{Cl}^- \rightarrow \text{H} \equiv \text{Cl}^- \)
- NH\(_3\): \( \text{N}^3+ \cdot \text{H}^+ \cdot \text{H}^+ \rightarrow \text{N} \equiv \text{H} \equiv \text{H} \)
- H\(_2\)S: \( \text{S}^2+ + \cdot \text{H}^+ + \cdot \text{H}^+ \rightarrow \text{S} \equiv \text{H} \equiv \text{H} \)
- O\(_2\): \( \text{O} \equiv \text{O} \)
- Br\(_2\): \( \text{Br} \equiv \text{Br} \)
- H\(_2\)O\(_2\): \( \text{H} \equiv \text{O} \equiv \text{O} \equiv \text{H} \)
- C\(_2\)H\(_4\): \( \text{H} \equiv \text{C} \equiv \text{C} \equiv \text{H} \)

### Chemical Compounds

- H\(_2\)S
- H\(_2\)O\(_2\)
- C\(_2\)H\(_4\)
- CH\(_3\)NH\(_2\)
14. (a) 
\[ \text{H} \quad \text{O} \quad \text{H} + \text{H}^+ \rightarrow \left[ \begin{array}{c} \text{H} \\ \text{O} \\ \text{H} \end{array} \right]^+ \]

(b) There are covalent and coordinate covalent bonds within the hydronium ion.
(c) Ionic bonds. (Note that it will also form hydrogen bonds.)

15. It is incorrect to show the structural formula of \( \text{H}_2\text{S} \) as \( \text{H}--\text{H}--\text{S} \). The central H atom is shown as bonding to two other atoms — yet the bonding capacity of hydrogen is one. The correct structural formula is \( \text{H}--\text{S}--\text{H} \) and its formation is shown below:
\[
\text{S}^--\text{H} : \text{H} + \text{H} : \text{H} \rightarrow \text{H} -- \text{S} -- \text{H} \rightarrow \text{H} \rightarrow \text{S} -- \text{H}
\]

SECTION 2.3 QUESTIONS

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Understanding Concepts
1. (a) Properties of compound A: a solid at SATP, is hard and brittle, has a high melting point, and its solution conducts electricity.
   Properties of compound B: may be a solid, a liquid, or a gas at SATP — if a solid it is soft, waxy or flexible, has a low melting point, and its solution does not conduct electricity.
(b) Compound A is an ionic compound involving the bonding of a metal with a nonmetal. Sodium, a metal, has formed an ionic bond with fluorine, a nonmetal.
   Compound B is a molecular compound involving the bonding of a nonmetal with another nonmetal. Nitrogen, a nonmetal, has formed covalent bonds with fluorine, also a nonmetal.
(c) \( \text{Na}^+ + \cdot\text{F}^- \rightarrow [\text{Na}]^+ + [\cdot\text{F}^-]^- \)

2. The difference in electronegativities between aluminum and oxygen is strong enough to transfer electrons from the metal atom to the nonmetal atom, producing ionic bonds and an ionic compound. Metals and nonmetals show a periodic trend to form ionic compounds, and the metals of Group 13 show a periodic trend to form ionic oxides (compounds composed of a metal and oxygen) when burned in air.
   The difference in electronegativities between carbon and oxygen is not strong enough to transfer electrons, and instead, the electrons are shared in a covalent bond, resulting in the formation of a molecular compound. Nonmetals show a periodic trend to combine with other nonmetals to form molecular compounds.

3. (a) Group 1 metals readily react with the nonmetal elements in Group 17 to form ionic compounds, with the general formula \( \text{MX} \).
   (b) The compound \( \text{MX} \) is a solid at SATP, is hard and brittle, has a high melting point, and its aqueous solution conducts electricity.

4. (a) \( \text{NaCl} \) is an ionic compound; \( \text{Cl}_2 \) is a covalent compound.
   (b) \( \text{NaCl} \) is a solid at SATP because of the strong ionic bonds. The simultaneous forces of attraction between the positive and negative ions hold the ions firmly in a rigid structure.
   \( \text{Cl}_2 \) is a gas at SATP because of the weak intermolecular forces between the \( \text{Cl}_2 \) molecules. The covalent bonds between the Cl atoms are strong, but with the addition of a relatively small amount of heat, the intermolecular forces between \( \text{Cl}_2 \) molecules are easily overcome, causing the molecular compound to be a gas at SATP.
5. Some of the molecular compounds that can be created using oxygen and sulfur include SO, SO$_2$, SO$_3$. 

\[
\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{S}}}} + \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}} \rightarrow \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{S}}} = \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}}.
\]

\[
\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{S}}} + \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}} + \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}} \rightarrow \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}} - \overset{\cdot}{\overset{\cdot}{S}} = \overset{\cdot}{\overset{\cdot}{O}}.
\]

\[
\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{S}}} + \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}} + \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}} + \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}} \rightarrow \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}} - \overset{\cdot}{\overset{\cdot}{S}} = \overset{\cdot}{\overset{\cdot}{O}}.
\]

From the examples above, SO contains only multiple bonds.

**Applying Inquiry Skills**

6. **Experimental Design**

**Solubility:** Obtain a small amount of compound A. Observe and record its state at the ambient temperature. Add a small quantity of the substance to about 10 mL of distilled water. Stir the mixture with a stirring rod and note whether the chemical dissolves. Many ionic compounds readily dissolve in water.

**Conductivity:** Obtain a small sample of distilled water in a beaker. Use a low-voltage conductivity apparatus to test the electrical conductivity of the sample. The apparatus should indicate a reading of zero. Test the electrical conductivity of the mixture from the above solubility procedure and record observations. Ionic compounds (many of which dissolve readily in water) form solutions that conduct electricity.

Repeat the procedure for compound B. Molecular compounds (some of which dissolve in water) form solutions that do not conduct electricity.

7. **Experimental Design**

**State, Hardness and Brittleness:** Obtain a small-sized piece of NaCl\(_{(s)}\) in road salt form. Observe and, in a table, record its state at SATP, its hardness, and brittleness.

**Solubility:** Pour about 10 mL of distilled water into a 50-mL beaker. Add a small quantity of the road salt to the water. Use a stirring rod to stir the mixture. Note whether the road salt dissolves.

**Conductivity:** Obtain a small sample of distilled water in a beaker. Test the electrical conductivity of the sample. The apparatus should indicate a reading of zero. Test the conductivity of the mixture of sodium chloride and water from the Solubility procedure. Record observations.

---

**2.4 ELECTRONEGATIVITY, POLAR BONDS, AND POLAR MOLECULES**

**PRACTICE**

(Page 84)

**Understanding Concepts**

1. To predict whether a chemical bond between two atoms will be ionic, polar covalent, or covalent, we must consider the electronegativities of the elements involved. The absolute value of the difference in electronegativities of two bonded atoms provides a measure of the polarity in the bond: the greater the difference, the more polar the bond. By convention, a difference in electronegativity greater than 1.7 indicates an ionic bond.

2. (a) covalent (f) covalent
   (b) covalent (g) covalent
   (c) ionic (h) ionic
   (d) ionic (i) covalent
   (e) covalent

Si and O would be the most polar of the covalent bonds.

3. (a) H—F (e) N—H
   (b) C—O (f) P—O
   (c) O—H (g) C—N
   (d) P—Cl
4. (a) $\text{H}_2\text{O}$

(b) $\text{Br}_2$

(c) $\text{HBr}$

(d) $\text{PCl}_3$

(e) $\text{OF}_2$

PRACTICE

Understanding Concepts

5. If a molecule that contains polar covalent bonds is quite symmetrical, it lacks oppositely charged ends and is not polar.

6. While the difference in electronegativity values between carbon and hydrogen results in polar covalent bonds (the difference is 0.4 for each carbon — hydrogen bond), the molecule of methane, $\text{CH}_4(\text{g})$, is quite symmetrical, and lacks oppositely charged ends. Therefore, it is not polar.

Try This Activity: Molecular Models

(a) $\text{O} \equiv \text{C} \equiv \text{O}$

(b) $\text{O} \equiv \text{C} \equiv \text{O}$

(c) Carbon dioxide required a double bond in order to complete its octet. Nitrogen required a triple bond in order to complete its octet.

(d) Water, hydrogen sulfide, hydrogen fluoride, and ammonia would be expected to be polar.
7. In ice, the hydrogen bonds between the molecules result in a regular hexagonal crystal structure that forms an open lattice with a great deal of empty space between the molecules. This causes ice to be less dense than liquid water. Thus, as water freezes, the ice floats to the top of the water surface, and as time passes, the lake freezes from the top down.

8. (a) polar (d) nonpolar
   (b) nonpolar (e) nonpolar
   (c) polar (f) nonpolar

9. $\text{PCl}_3(s)$, $\text{H}_2\text{C}_2\text{H}_3\text{O}_2(\text{aq})$, $\text{CCl}_4(l)$ were classified as nonpolar but contain polar covalent bonds. Due to the symmetrical shape of these molecules, they lack oppositely charged ends and are therefore not polar.

10. A dipole–dipole force describes in general the attractive force acting between polar molecules. A hydrogen bond is a relatively strong dipole–dipole force acting specifically between a positive hydrogen atom of one molecule and a highly electronegative atom (F, O, or N) in another molecule.

11. (a) London dispersion force
    (b) Hydrogen bond
    (c) Hydrogen bond

12. (a) $\text{I}_2(s)$ is a nonpolar molecule. Thus it is the London dispersion force that exists between all molecules — both polar and nonpolar — that is responsible for the intermolecular attraction between molecules of $\text{I}_2(s)$.
    (b) $\text{H}_2\text{O}(l)$ is a highly polar molecule containing an O–H bond. Thus it is the hydrogen bond force — a relatively strong dipole–dipole force between a positive hydrogen atom of one molecule and a highly electronegative atom (F, O, or N) in another molecule — that is responsible for the intermolecular attraction between molecules of $\text{H}_2\text{O}(l)$.
    (c) $\text{NH}_3(g)$ is a highly polar molecule containing an N–H bond, and so forms hydrogen bonds with other molecules.

**Applying Inquiry Skills**

13. **Prediction**
    (a) The molecules that will be affected by the charged object are $\text{NCl}_3$, $\text{H}_2\text{O}$, $\text{CH}_3\text{OH}$, and $\text{H}_2\text{O}_2$. The charged object will have an effect on polar molecules (only) because the polar molecules are slightly charged at each end.

**Evidence**
(b) Samples 1 to 3 are $\text{Br}_2$, $\text{CCl}_4$, and vegetable oil. Samples 4 to 7 are $\text{NCl}_3$, $\text{H}_2\text{O}$, $\text{CH}_3\text{OH}$, and $\text{H}_2\text{O}_2$.

**Analysis**
(c) The evidence shows that a charged object will have an effect on a thin stream of liquids that are composed of polar molecules, but will have no effect on liquids composed of nonpolar molecules. The thin stream of liquids composed of polar molecules will be attracted to the charged object. Of the seven sample liquids, only $\text{NCl}_3$, $\text{H}_2\text{O}$, $\text{CH}_3\text{OH}$, and $\text{H}_2\text{O}_2$ are polar, and therefore affected by the charged object.

**Synthesis**
(d) The polar molecules — $\text{NCl}_3$, $\text{H}_2\text{O}$, $\text{CH}_3\text{OH}$, and $\text{H}_2\text{O}_2$ — are positively charged at one end, and negatively charged at the other because of electronegativity differences. Thus, the end of the polar molecule that has the opposite charge of the charged object will be attracted to the charged object. This is similar to dipole–dipole forces.

(e) The liquids were affected by both positive and negative charges because the polar molecules of the liquid are positively charged at one end, and negatively charged at the other. If a positively charged object is used, the negatively charged end of the polar molecule will be attracted to the charged object, and if a negatively charged object is used, the positively charged end of the polar molecule will be attracted to the charged object.
SECTION 2.4 QUESTIONS

(PAGE 89)

Understanding Concepts
1. Covalent bonds and ionic bonds are the forces that bond atoms and ions together within a compound — the intramolecular forces. These forces are sufficient to explain the existence of molecular and ionic compounds, and to explain many of the properties of ionic compounds, but they aren’t sufficient to explain the physical state of molecular compounds. If covalent bonds were the only forces at work, molecular compounds would all be gases, as there would be no attraction between the molecules strong enough to order the molecules into solids or liquids. The concepts of the polar molecule and small charges on atoms that result in intermolecular forces help to explain why these molecular compounds are not all gases at SATP.

2. When the atoms are identical, such as in a chlorine molecule, the electrons are shared equally. However, this is not the case for a compound like hydrogen chloride, where electrons are shared between two different elements. In this situation, the sharing is unequal, as the bonding electrons in the H—Cl bond spend more time near the chlorine atom than near the hydrogen atom. This is because of chlorine’s greater attraction for electrons.

3. (a) Both BCl₃ and NCl₃ are molecular compounds. By convention, compounds with bonds that have electronegativity differences less than or equal to 1.7, have covalent-type bonds and are classified as molecular compounds.

   (b) The bonds between B—Cl and N—Cl are similar in that they are both covalent-type bonds and involve the sharing of a pair of electrons. The bonds are different in that the B—Cl bonds are polar covalent bonds due to differences in electronegativities. In this situation, the sharing of electrons is unequal, as the bonding electrons in the B—Cl bond spend more time near the chlorine atom than near the boron atom. This is because of chlorine’s greater attraction for electrons. Another difference is that the boron atom in BCl₃ does not have an octet in its valence shell.

   (c) BCl₃ is quite symmetrical and lacks oppositely charged ends. Thus it is not a polar molecule. Intermolecular attractions would be due to the weaker London dispersion forces. NCl₃ is a polar molecule due to having nitrogen at one end. Intermolecular attractions would be due to the stronger dipole-dipole forces.

4. Since a molecule of carbon tetrachloride, CCl₄, is quite symmetrical, it lacks oppositely charged ends and is not polar.

2.5 THE NAMES AND FORMULAS OF COMPOUNDS

(PAGE 93)

Understanding Concepts
1. Substances were named in a variety of ways. In some cases, the name referred to the use of the compound; in other cases, it incorporated an obvious property, or perhaps referred to the sources of the substance.

2. (a) muriatic acid
   (b) baking soda
   (c) laughing gas
   (d) grain alcohol

3. A binary compound is composed of two kinds of elements.

4. In the formula of a binary ionic compound, the metal cation is always written first, followed by the nonmetal anion. The name of the metal is stated in full and the name of the nonmetal ion has an -ide suffix; for example, NaCl(s) is sodium chloride.
5. Most transition metals and some representative metals can form more than one kind of ion. Metals that can have more than one valence, or charge, are classified as multivalent. For example, iron can form an Fe$^{2+}$ ion or an Fe$^{3+}$ ion.

6. (a) CaF$_2$  
(b) Na$_2$S  
(c) AlN  
(d) AlCl$_3$  
(e) K$_2$O  
(f) CaCl$_2$  
(g) CuS  
(h) PbBr$_2$  
(i) AgI  
(j) Ba$_3$N$_2$  
(k) FeF$_2$

7. (a) sodium chloride  
(b) calcium oxide  
(c) calcium chloride  
(d) magnesium oxide  
(e) aluminum oxide  
(f) zinc sulfide  
(g) tin(IV) oxide  
(h) copper(I) sulfide  
(i) lead(IV) sulfide  
(j) iron(III) oxide  
(k) molybdenum oxide  
(l) silver sulfide  
(m) zinc oxide  
(n) tin(IV) oxide

8. (a) sodium oxide  
(b) tin(IV) chloride  
(c) zinc iodide  
(d) strontium chloride  
(e) aluminum bromide  
(f) lead(IV) chloride  
(g) nickel(III) oxide  
(h) silver sulfide  
(i) iron(II) chloride  
(j) potassium bromide  
(k) copper(II) iodide  
(l) nickel(II) sulfide

9. (a) SrO - strontium oxide  
(b) Na$_2$S - sodium sulfide  
(c) AgI - silver iodide  
(d) BaF$_2$ - barium fluoride  
(e) CaBr$_2$ - calcium bromide  
(f) LiCl - lithium chloride

10. (a) HgS  
(b) MoS$_2$  
(c) MnO$_2$  
(d) NiBr$_2$  
(e) CuCl$_2$  
(f) FeI$_3$

11. (a) iron(II) sulfide  
(b) lead(IV) bromide  
(c) tin(II) chloride

Reflecting
12. Because old systems die hard. For example, the -ous and -ic suffixes are still used extensively in industry.

PRACTICE
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Understanding Concepts
13. Tertiary compounds are composed of three different elements.
14. (a) Ionic compounds can be composed of a single ion and a polyatomic ion.  
(b) Some ionic compounds are composed of a single ion and a polyatomic ion that includes oxygen. A polyatomic ion that includes oxygen is called an oxyanion.  
(c) Ionic compounds that form crystals that contain molecules of water within the crystal structure are referred to as hydrates.
15. (a) sodium nitrate  
(b) sodium nitrite  
(c) copper(II) nitrate  
(d) copper(I) nitrate  
(e) aluminum sulfate  
(f) calcium hydroxide  
(g) lead(II) carbonate  
(h) tin(II) phosphate  
(i) iron(III) sulfate
16. (a) CaCO₃   (f) (NH₄)₃PO₄  
(b) NaHCO₃   (g) CuSO₄  
(c) NaClO   (h) NaOH  
(d) CaSO₄   (i) KmNO₄  
(e) NH₄NO₃  
17. (a) lithium chlorate   (n) silver sulfate  
(b) barium sulfate   (o) mercury(II) bromate  
(c) mercury(I) carbonate   (p) iron(III) carbonate  
(d) magnesium nitrate   (q) ammonium hypochlorate  
(e) iron(III) bromate   (r) gold(III) nitrate  
(f) sodium phosphate   (s) magnesium bromate  
(g) ammonium iodate   (t) sodium iodate  
(h) gold(I) acetate   (u) zinc chlorite  
(i) zinc phosphate   (v) tin(II) carbonate  
(j) antimony(V) chloride   (w) strontium sulfate  
(k) manganese(II) sulfite   (x) nickel(III) phosphate  
(l) potassium hypobromite   (y) copper(II) acetate  
(m) aluminum perphosphate   (z) barium perphosphate  
18. (a) copper(I) hypophosphite   CuPO₂  
(b) tin(IV) chlorite   Sn(ClO₂)₄  
(c) iron(II) bromate   Fe(BrO₃)₂  
(d) iron(III) chlorite   Fe(ClO₂)₃  
(e) lead(IV) sulfate   Pb(SO₄)₂  
19. (a) copper(II) pentahydrate  
(b) sodium sulfate decahydrate  
(c) magnesium sulfate heptahydrate  
20. (a) Fe₂O₃·3H₂O   (d) Cd(NO₃)₂·4H₂O  
(b) AlCl₃·6H₂O   (e) LiCl·4H₂O  
(c) NaS₂O₃·5H₂O   (f) CaCl₂·2H₂O  
21. When heat is applied to a hydrate, it will decompose to produce water vapour and an associated ionic compound. When this water, called water of hydration, is removed, the product is referred to as anhydrous.  

PRACTICE  

(Page 98)  

Understanding Concepts  

22. (a) N₂   (n) SF₄  
(b) CO₂   (o) PCl₅  
(c) CO⁻   (p) S₂Cl₂  
(d) NO₂   (q) CCl₄  
(e) NO⁺   (r) SO₃  
(f) N₂O   (s) SF₆  
(g) N₂O₄   (t) ClO₂  
(h) SO₃₂⁻   (u) N₂O₅  
(i) I₂O₅   (v) PCl₃  
(j) SiF₄   (w) SiCl₄  
(k) BF₃   (x) CS₂  
(l) PI₃   (y) PbBr₃  
(m) P₂O₅   (z) CF₄  
23. (a) sulfur hexafluoride   (f) iodine heptafluoride  
(b) dinitrogen trioxide   (g) boron trifluoride  
(c) nitrogen dioxide   (h) diphosphorus pentasulfide  
(d) phosphorus trichloride   (i) diphosphorus pentoxide  
(e) phosphorus pentachloride  

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Chapter 2 Chemical Bonding  

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Understanding Concepts

24. (a) $\text{HCl}_{(aq)}$  
(b) $\text{HCl}_{(aq)}$  
(c) $\text{H}_2\text{SO}_4_{(aq)}$  
(d) $\text{H}_2\text{SO}_4_{(aq)}$  
(e) $\text{HC}_2\text{H}_3\text{O}_2_{(aq)}$  
(f) $\text{HC}_2\text{H}_3\text{O}_2_{(aq)}$  
(g) $\text{HNO}_2_{(aq)}$  
(h) $\text{HNO}_3_{(aq)}$  
(i) $\text{HBr}_{(aq)}$  
(j) $\text{H}_2\text{SO}_2_{(aq)}$  
(k) $\text{HI}_{(aq)}$  
(l) $\text{HClO}_4_{(aq)}$

25. Classical IUPAC  
(a) sulfurous acid  
(b) phosphoric acid  
(c) hydrocyanic acid  
(d) carbonic acid  
(e) hydrosulfuric acid  
(f) hydrochloric acid  
(g) hydrosulfuric acid  
(h) sulfuric acid  
(i) phosphoric acid

SECTION 2.5 QUESTIONS

Understanding Concepts

1. (a) $\text{NH}_3$, $\text{HCN}$  
(b) Ammonia and hydrogen cyanide are classified as covalent molecules.  
(c) Hydrogen cyanide is a polar covalent molecule that ionizes in water to form $\text{H}^+$ and $\text{CN}^-$. The ionic nature of the compound could be verified by dissolving the substance in water and testing for electrical conductivity. The covalent nature of the compound could be verified by calculating the electronegativity difference between $\text{H}$ and $\text{C}$, and between $\text{C}$ and $\text{N}$. The differences are not greater than 1.7.  
(d) If hydrogen cyanide is added to water, aqueous hydrogen cyanide, also known as hydrocyanic acid, is formed. The substance might cause blue litmus paper to turn red, indicating an acidic solution. (Actually, hydrocyanic acid is a very weak acid and may not turn blue litmus paper red.)

2. (a) $\text{KOH}$ is an ionic hydroxide, and its aqueous solution is a base. The bond consists of a $\text{K}^+$ ion, and an $\text{OH}^{-}$ ion. Chemists have discovered that all aqueous solutions of ionic hydroxides are bases. $\text{HCl}$ as a gas is covalent, and its aqueous solution is an acid. The bond consists of an unequal sharing of a pair of electrons. When dissolved in water, the resulting aqueous solution displays a set of specific properties called acidic.  
(b) $[\text{K}^+ \quad [\text{Cl}^-\text{H}^-] \rightarrow \text{KOH}$  
(c) aqueous potassium hydroxide, aqueous hydrogen chloride  
(d) Bases are reactive — for example, bases react with proteins to break them down into smaller molecules. Precautions must be taken when handling bases. Acids are reactive and can combine with many other substances. Acids must be treated with care, as they can be very corrosive and can cause serious damage to the environment.

3. The student is to write the IUPAC names and formulas for as many compounds as possible, using only the following elements: $\text{K}$, $\text{C}$, $\text{H}$, $\text{F}$, $\text{Mg}$, $\text{O}$, $\text{Cl}$, and $\text{Na}$. Use $\text{K}$ as an example:  
(a) $\text{KCl}$, $\text{KOH}$, $\text{KClO}_3$
(b) KCl — potassium chloride, KOH — potassium hydroxide, KClO₃ — potassium chlorate.

(c) KCl — ionic, KOH — ionic, KClO₃ — ionic.

\[
\text{[K]}^+ \text{[Cl]}^- \rightarrow \text{KCl}
\]

\[
\text{[K]}^+ \text{[:\ddot{O} --- H]}^- \rightarrow \text{KOH}
\]

\[
\text{[K]}^+ \text{[:\ddot{O} --- \ddot{O}]} \rightarrow \text{KClO₃}
\]

(d) KCl — binary, KOH — tertiary and basic, KClO₃ — tertiary.

(e) KCl — ionic bonds only, KOH — ionic and covalent bonds, KClO₃ — ionic and covalent bonds.

## CHAPTER 2 SUMMARY

**Make a Summary**

(Page 102)

The examples of NaCl and H₂O are used in the table below. The student is to include as many examples of compounds as possible for each type of intramolecular bond.

### Table 1: Summarizing Bonds and Forces

<table>
<thead>
<tr>
<th>Compound</th>
<th>Properties</th>
<th>Electron dot diagram/ Lewis structure</th>
<th>Intramolecular bond type</th>
<th>Polarity</th>
<th>Intermolecular forces</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCl</td>
<td>Solid at SATP, hard and brittle, high melting point, its solution conducts electricity.</td>
<td>[Na]⁺ [Cl]⁻</td>
<td>Ionic</td>
<td>Ionic</td>
<td>Locked in a regular structure, held by the balance of attractive bonds and electrical repulsion.</td>
</tr>
<tr>
<td>H₂O</td>
<td>Liquid at SATP, low boiling point.</td>
<td>H—O—H or H(\dddot{O}---\dddot{H})</td>
<td>Covalent</td>
<td>Polar</td>
<td>Hydrogen bonds.</td>
</tr>
</tbody>
</table>

**Reflect on your Learning**

(Page 102)

By the end of the chapter the student should have developed a more in-depth understanding of why atoms form compounds, an awareness of the many different compounds that are possible, the types of forces present between atoms in compounds, and how the forces that hold atoms together in a compound determine the chemical properties of the compound.

## CHAPTER 2 REVIEW

(Page 103)

**Understanding Concepts**

1. When elements that are found in the “metals” position in the periodic table react with elements that are found in the “nonmetals” position in the periodic table, they form ionic compounds that have ionic bonds. When elements that are found in the “nonmetals” position in the periodic table react with elements also found in the “nonmetals” position in the periodic table, they form molecular compounds that have covalent bonds.
2. An ionic bond occurs when one or more valence electrons are transferred from a metal atom to a nonmetal atom. This leaves the metal atom as a positive ion, or cation, and the nonmetal atom as a negative ion, or anion. An ionic bond is the electrostatic attraction between positive and negative ions in a compound.

A covalent bond arises from the simultaneous attraction of two nuclei for a shared pair of electrons. The result is a covalent bond—a shared pair of electrons held between two nonmetal atoms that hold the atoms together in a molecule.

3. (a) The properties of ionic compounds: Are solid at SATP, with hard surfaces, are brittle, have high melting points, and form solutions that conduct electricity. The properties are due to the strong ionic bonds, simultaneous forces of attraction between the positive and negative ions, which hold the ions firmly in a rigid structure. The solid state, hardness, brittleness, and the high melting point result from the strong attractions, which occur in the crystal structure. And because the ionic bonds break down in water, the resulting ions are free to move in solution and conduct electricity.

   The properties of molecular compounds: May be solids, liquids, or gases at SATP, and are soft, waxy, or flexible. Covalent bonds between the atoms are strong. However, the intermolecular forces in molecular compounds are weaker in comparison—adding a relatively small amount of heat will cause a solid molecular compound to change state from a solid to a liquid, and then to a gas.

(b) Ionic compounds (many of which dissolve readily in water) form solutions that conduct electricity. Because the ionic bonds often break down in water, the resulting ions are free to move in solution and conduct electricity. Molecular compounds form solutions that do not generally conduct electricity.

4. (a) Intramolecular bonding is the force that bonds atoms and ions together in a compound. One main type of intramolecular bonding is ionic bonding. An example of ionic bonding is the electrostatic attraction that occurs between the $[\text{Na}]^+$ cation and the $[\text{Cl}]^-\text{' anion to form the ionic compound NaCl. Another main type of intramolecular bonding is covalent bonding. An example of covalent bonding is the sharing of a pair of electrons that occurs between hydrogen and chlorine to form the molecular compound HCl.

(b) Ionic bonding results in compounds that are solid at SATP with hard surfaces, are brittle, have high melting points, and form solutions that conduct electricity. Covalent bonding results in compounds that may be solids, liquids, or gases at SATP, and are soft, waxy, or flexible. Covalently bonded compounds form solutions that do not generally conduct electricity.

5. (a) two nonmetal atoms that are sharing a pair of electrons
(b) oppositely charged ends of polar molecules
(c) a positive hydrogen atom of one molecule and a highly electronegative atom (F, O, or N) in another molecule
(d) a positively charged ion (cation) of a metal and a negatively charged ion (anion) of a nonmetal

6. The chemical formulas of ionic compounds consist of a metal joined to a nonmetal. Examples are NaCl, CuSO4, and NaHCO3. The chemical formulas of molecular compounds consist of nonmetals combined with other nonmetals. Examples are SO2, CO2, and NH3.

7. Halogens tend to form diatomic molecules because they have only one bonding electron, and thus a capacity to bond with only one other atom.

8. Students will reproduce the bonding continuum of Figure 3 on p. 84. Cl2, with difference in electronegativities of 0, will be placed at the far right (covalent). NaCl, with an electronegativity difference of 2.1, should be placed left of centre, in the “ionic” area. Na–Cl involves an electron transfer, resulting in the formation of cations and anions that are attracted to each other. Cl–Cl involves equal sharing of a pair of electrons.

9. (a) Ca \(\cdot\text{Ca}\cdot\)
(b) Al \(\cdot\text{Al}\cdot\)
(c) K \(\cdot\text{K}\cdot\)
(d) N \(\cdot\text{N}\cdot\)
(e) S \(\cdot\text{S}\cdot\)
(f) Br \(\cdot\text{Br}\cdot\)
(g) Ne \(\cdot\text{Ne}\cdot\)

10. (a) covalent
(b) polar covalent
(c) polar covalent
(d) ionic
(e) ionic
(f) ionic

11. (a) $\text{Na}_2\text{O}_{(s)}$ is ionic, $\text{MgO}_{(s)}$ is ionic, $\text{Al}_2\text{O}_3(s)$ is ionic, $\text{SiO}_2(s)$ is molecular, $\text{P}_2\text{O}_5(s)$ is molecular, $\text{SO}_2(g)$ is molecular, and $\text{Cl}_2\text{O}_3(g)$ is molecular.
2. The difference in electronegativity of the constituent elements determines whether the compound has ionic bonds or covalent-type bonds. An electronegativity difference greater than 1.7 indicates an ionic bond. Ionic bonds result in ionic properties, and covalent-type bonds result in molecular properties.

### Table 1: Structures of Covalent Compounds

<table>
<thead>
<tr>
<th>Compound</th>
<th>Lewis Structure</th>
<th>Structural Formula</th>
<th>Types of Bonds</th>
</tr>
</thead>
<tbody>
<tr>
<td>HF(_{(g)})</td>
<td>(\cdot\text{H} - \cdot\text{F})</td>
<td>(\text{H} - \text{F})</td>
<td>Polar covalent</td>
</tr>
<tr>
<td>BCl(_3(g))</td>
<td>(\cdot\text{Cl} : \cdot\text{Cl} : \cdot\text{Cl})</td>
<td>(\text{Cl} - \text{B} - \text{Cl})</td>
<td>Polar covalent</td>
</tr>
<tr>
<td>SiH(_4(g))</td>
<td>(\cdot\text{H} - \cdot\text{Si} - \cdot\text{H} - \cdot\text{H})</td>
<td>(\text{H} - \text{Si} - \text{H})</td>
<td>Polar covalent</td>
</tr>
<tr>
<td>CCl(_4(l))</td>
<td>(\cdot\text{Cl} : \cdot\text{Cl} : \cdot\text{Cl} : \cdot\text{Cl})</td>
<td>(\text{Cl} - \text{C} - \text{Cl})</td>
<td>Polar covalent</td>
</tr>
<tr>
<td>NCl(_3(g))</td>
<td>(\cdot\text{Cl} : \cdot\text{N} : \cdot\text{Cl})</td>
<td>(\text{Cl} - \text{N} - \text{Cl})</td>
<td>Polar covalent</td>
</tr>
<tr>
<td>H(_2)O(_2(l))</td>
<td>(\cdot\text{H} - \cdot\text{O} - \cdot\text{O} - \cdot\text{H})</td>
<td>(\text{H} - \text{O} - \text{O} - \text{H})</td>
<td>Polar covalent and covalent</td>
</tr>
<tr>
<td>CO(_2(g))</td>
<td>(\cdot\text{O} = \cdot\text{C} = \cdot\text{O})</td>
<td>(\text{O} = \text{C} = \text{O})</td>
<td>Polar covalent</td>
</tr>
<tr>
<td>HCN(_(g))</td>
<td>(\cdot\text{H} - \cdot\text{C} \equiv \cdot\text{N})</td>
<td>(\text{H} - \text{C} \equiv \text{N})</td>
<td>Polar covalent</td>
</tr>
</tbody>
</table>
13. (a) Cl is most electronegative and will have a $\delta^-$ charge.

\[ \cdot \hat{\text{Cl}} + \cdot \hat{\text{Cl}} \rightarrow \cdot \hat{\text{Cl}}^- \cdot \hat{\text{Cl}}^+ \]

(b) F is most electronegative and will have a $\delta^-$ charge.

\[ \cdot \hat{\text{Ca}} + \cdot \hat{\text{F}} \rightarrow [\text{Ca}^+] [\cdot \hat{\text{F}}^2]^- \]

(c) Cl is most electronegative and will have a $\delta^-$ charge.

\[ \cdot \hat{\text{Al}} + \cdot \hat{\text{Cl}} \rightarrow [\cdot \hat{\text{Al}}^+] [\cdot \hat{\text{Cl}}]^2^- \]

(d) O is most electronegative and will have a $\delta^-$ charge.

\[ \cdot \hat{\text{Si}} + \cdot \hat{\text{O}} \rightarrow \cdot \hat{\text{Si}}^+ = \cdot \hat{\text{O}}^- \]

(e) O is most electronegative and will have a $\delta^-$ charge.

\[ \cdot \hat{\text{C}} + \cdot \hat{\text{O}} \rightarrow \cdot \hat{\text{C}}^+ = \cdot \hat{\text{O}}^- \]

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

\[
\begin{array}{c}
\cdot \hat{\text{Cl}}
\end{array}
\]

\[ \text{NH}_4\text{Cl}(s) \text{ contains a coordinate covalent bond. A hydrogen ion, which has no electrons of its own, has bonded to an NH}_4^{+} \text{ molecule by sharing the unbonded, lone pair of electrons in NH}_3^{(g)} \cdot \]

(b) \[\begin{array}{c}
\cdot \hat{\text{F}} \\
\cdot \hat{\text{B}} \\
\cdot \hat{\text{F}}
\end{array}
\]

\[ \text{HF}_3\text{BF}_3(g) \text{ contains a coordinate covalent bond. The boron in BF}_3(g) \text{, which does not have an octet in its valence shell, has bonded to NH}_3 \text{ by sharing the unbonded, lone pair of electrons in NH}_3. \]

15. (a) NaHSO$_4$(s) (h) NaNO$_3$(s)
(b) NaOH$_3$(s) (i) H$_2$PO$_4$(aq)
(c) CO$_2$(g) (j) I$_2$(s)
(d) HCO$_2$(aq) (k) Al$_2$O$_3$(s)
(e) NaS$_2$O$_3$ : 5 H$_2$O$_3$(s) (l) KOH$_3$(s)
(f) NaClO$_3$(s) (m) HCO$_3$(aq)
(g) S$_8$(s)

16. (a) MgBr$_2$ (n) H$_2$CO$_3$(aq)
(b) CS$_2$ (o) Ca(OH)$_2$
(c) Hg(NO$_3$)$_2$ (p) Zn(ClO)$_3$
(d) HCl$_3$(aq) (q) Pb(ClO)$_3$(aq)
(e) LiOH$_3$ (r) PBr$_5$
(f) $\text{Ag}_2\text{CO}_3$  (s) $\text{AsCl}_5$
(g) $\text{Al(ClO}_4)_3$  (t) $\text{Bi(NO}_3)_3$
(h) $\text{CuSO}_4$  (u) $\text{NaClO}$
(i) $\text{SO}_3$  (v) $\text{OCl}_2$
(j) $\text{NiPO}_4$  (w) $\text{SnBr}_2$
(k) $\text{MgO}$  (x) $\text{H}_2\text{SO}_4(\text{aq})$
(l) $\text{N}_2\text{O}$  (y) $\text{KOH}$
(m) $\text{FeSO}_5$  (z) $\text{BaCO}_3$

17. (a) $\text{NH}_4\text{H}_2\text{PO}_3$
(b) $\text{LiHSO}_3$
(c) $\text{KHSO}_4$
(d) $\text{BaCl}_2 \cdot 3 \text{H}_2\text{O}$
(e) $\text{NaH}_2\text{PO}_4$
(f) $\text{NaHCO}_3$

18. (a) calcium carbonate
(b) diphosphorus pentoxide
(c) magnesium sulfate heptahydrate
(d) dinitrogen monoxide
(e) sodium silicate
(f) calcium hydrogen carbonate
(g) hydrochloric acid
(h) copper(II) sulfate pentahydrate
(i) sulphuric acid
(j) calcium hydroxide
(k) sulfur trioxide
(l) sodium fluoride

19. (a) sodium chloride
(b) diphosphorus trioxide
(c) nitric acid
(d) lead(II) acetate
(e) ammonium hypochlorite
(f) tin(IV) bromate
(g) antimony(III) oxide
(h) zinc iodate
(i) iron(II) pernitrate
(j) calcium hydroxide
(k) potassium iodide
(l) sulfur difluoride
(m) hydrobromic acid
(n) copper(II) carbonate
(o) aluminum sulfite
(p) ammonium hydroxide
(q) barium acetate
(r) iodine monochloride
(s) gold(III) chloride
(t) magnesium sulfide
(u) dinitrogen difluoride
(v) nickel(II) sulfate
(w) hydrosulfuric acid
(x) silver bromate
(y) lithium perchlorate

20. (a) calcium hydrogen phosphate
(b) copper(II) sulfate heptahydrate
(c) sodium hydrogen phosphate
(d) lithium hydrogen carbonate
(e) potassium hydrogen sulfide

21. (a) potassium bromide $\text{KBr}_s$  $\text{K}^+ [\text{Br}^-]$
(b) silver iodide $\text{AgI}_s$  $\text{Ag}^+ [\text{I}^-]$
(c) lead(II) oxide $\text{PbO}_s$  $\text{Pb}^{2+} [\text{O}^{2-}]$
(d) zinc sulfide $\text{ZnS}_s$  $\text{Zn}^{2+} [\text{S}^{2-}]$
(e) copper(II) oxide $\text{CuO}_s$  $\text{Cu}^{2+} [\text{O}^{2-}]$
(f) lithium nitride $\text{Li}_3\text{N}_s$  $\text{Li}^{2+} [\text{N}^{3-}]$
Applying Inquiry Skills

22. (a) Analysis
Substance 1 is an ionic compound: KCl\(_{(aq)}\)
Substance 2 is an acid: HCl\(_{(aq)}\)
Substance 3 is soluble but not ionic: C\(_2\)H\(_5\)OH\(_{(aq)}\)
Substance 4 is a base: Ba(OH)\(_2\)\(_{(aq)}\)

(b) The water is used as a control. The result of the conductivity and litmus tests is the dependent variable. The different substances that are dissolved in water are the independent variables. Since water is used to prepare the solutions for each of the different substances, any change in the dependent variable from the control results can be attributed to the changing independent variable.

(c) Solutions 1, 2, and 4 all have high conductivities, and could have been involved in somebody getting electrocuted; however, the most likely to have been involved is KCl. If the solutions were of high concentration, it seems possible that the lawsuit would have been about corrosion rather than electrocution if the substance involved were the acid or the base. When KCl, HCl, and Ba(OH)\(_2\) are dissolved in water, the positive and negative ions dissociate. The ions can then carry an electrical current. The electronegativity differences between the constituent atoms of C\(_2\)H\(_5\)OH result in polar covalent bonds — C\(_2\)H\(_5\)OH is a polar molecule. With C\(_2\)H\(_5\)OH there are no positive and negative ions to dissociate, so the solution does not conduct electricity and could not have been involved in an electrocution.

\[ \text{K}^+ \quad [\cdot \text{Cl}]^- \]
\[ \text{H}^+ \quad [\cdot \text{Cl}]^- \]
\[ \text{Ba}^{2+} \quad 2[\text{OH}]^- \]
\[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array}
\quad \begin{array}{c}
\text{C} \\
\text{C} \\
\text{O} \\
\text{H} \\
\end{array}
\quad \begin{array}{c}
\text{H} \\
\text{H} \\
\end{array}
\]

Making Connections

23. (a) The student is to use the Internet to determine the IUPAC name and chemical formula for sal volatile. Ammonium carbonate: (NH\(_4\))\(_2\)CO\(_3\)

(b) The student is to predict the properties of the substance based upon the nature of the bonds. It is an ionic compound of the NH\(_4\)^+ and CO\(_3\)\(^{2-}\) ions. Students would predict that it is a solid, white, hard, and is highly soluble in water (hydrogen bonding). It is actually volatile at room temperature and even more so upon slight heating, releasing ammonia, water, and carbon dioxide.

(c) The student is to explain how the properties of the substance relate to the strong, sharp, ammonia smell. The release of ammonia (NH\(_3\)(g)) is responsible for the odour.

(d) The student is to use the Internet to find out how smelling salts used to be administered, how they worked, and comment on the safety concerns that this use would raise today. A bottle containing the salt, or the salt in an ammonia–water solution, was waved under the nose of the unconscious person. Human noses are extremely sensitive to ammonia, because it is relatively common in nature and also toxic. If our noses detect the smell of toxic chemicals, this information, vitally important for survival, can bypass the temporary shutdown that is characteristic of some forms of unconsciousness and “shock” the victim into consciousness. Ammonia dissolves readily in water, forming a basic solution that could damage the lining of the respiratory tract and the lungs. (Unconsciousness should also be respected. It is a defence of the body — shutting down consciousness prevents the brain, in an addled state caused by oxygen or nutrient deprivation or severe bodily trauma, from instructing the body to do something that endangers survival (like moving, for example). It also reduces the metabolic requirements of the brain in a time of shortage, preventing cell death.

\[ \text{H} \quad \text{GO TO} \quad \text{www.science.nelson.com, Chemistry 11, Teacher Centre.} \]

24. The student is to use the Internet to research the structure of boron nitride. Crystals with a hexagonal, graphite lattice are the most common form of boron nitride.

\[ \text{H} \quad \text{GO TO} \quad \text{www.science.nelson.com, Chemistry 11, Teacher Centre.} \]
25. The student is to use the Internet to research the foods that use glycerol, the physical and chemical properties of glycerol, the chemical formula and molecular shape of glycerol, and to use the knowledge of the molecular shape and intra- and intermolecular bonding to explain at least three of the properties of glycerol. The student is also to write a short report, outlining why glycerol is so useful to the food industry.
(a) In addition to the uses mentioned in the question, it is added to tobacco to keep it moist, and also to fruits, for the same purpose.
(b) Glycerol is also known as propanetriol (IUPAC), glycyl alcohol, and glycerine. At SATP it is a viscous liquid with a sweet taste. It dissolves readily in water (antifreeze). It reacts with a rather large range of organic substances, and with acids to produce esters.
(c) \( \text{CH}_2\text{OHCHOHCH}_2\text{OH} \); the molecule is polar and forms hydrogen bonds (explaining its solubility in water and its usefulness as a moistening agent, and also its viscosity in pure form).

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Exploring
26. The student is to use the Internet to research Fritz London’s work on intermolecular forces, and to write a brief report on his major findings. Dr. London (1900–1954) was born in what is now Wroclaw in Poland but was at that time Breslau in Germany. He emigrated to the United States in 1939 and became a citizen in 1945. A physicist and theoretical chemist, he worked on superconductivity, superfluids, and quantum chemistry (and is, of course, the London after whom London forces were named).

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