Bonding
Part 2

Atoms

Sharing of electrons

Molecule
Covalent bond

Transfer of electron

Positive ion
Negative ion

Ionic bond

Figure 2.6 Essential Cell Biology, 2/e, © 2001 Oxford Science

Name __________________________

Teacher _________________________

Regents Chemistry
Covalent Bonds

Definition: ________________________________

________________________________________

Types of Covalent Bonds

A. Nonpolar Covalent: _______________________

________________________________________

B. Polar Covalent ____________________________

________________________________________
C. Coordinate Covalent

Properties of Covalent Bonds
What is the duet rule? Which atom does this apply to?

What is the octet rule?

Draw the electron dot (Lewis) structures for each covalent molecule.

- F2
- O2
- H2S
- CO2
- HCl
- NH3
- H2
- N2
- PCl3
- SiH4
- BeCl2
- BH3

Which molecules above are exceptions to the octet rule?
Covalent bonding occurs when two or more nonmetals share electrons, attempting to attain a stable octet of electrons at least part of the time. For example:

\[ \text{H}^+ + \text{Cl}^- \rightarrow \text{HCl} \]

Note that hydrogen is content with 2, not 8, electrons.

Show how covalent bonding occurs in each of the following pairs of atoms. Atoms may share one, two or three pairs of electrons.

1. H + H (H₂)
2. F + F (F₂)
3. O + O (O₂)
4. N + N (N₂)
5. C + O (CO₂)
6. H + O (H₂O)
Electron-Dot Structures - Covalent Compounds

1. CH$_3$Cl

2. CH$_3$OH

3. H$_2$Te

4. OF$_2$

5. H$_2$S

6. PCl$_3$

7. SiO$_2$

8. CO$_2$

9. N$_2$

10. NH$_3$

11. HCN

12. HClO

13. C$_2$H$_4$

14. C$_2$H$_2$
Covalent Bonds

Covalent bonds are bonds formed by sharing electrons. The electrons of one atom are attracted to the protons of another, but neither atom pulls strongly enough to remove an electron from the other. Covalent bonds form when the electronegativity difference between the elements is less than 1.7 (see the Electronegativity table on the back of the Periodic Table) or when hydrogen behaves like a metal. When a covalent bond forms, no valence electrons are transferred, rather, they are shared. If the electronegativity difference is zero, the electrons are shared equally and the bond is nonpolar. If the electronegativity difference is greater than 0.4 but less than 1.7, the electrons are displaced towards the more electronegative element (nonmetal) and the bond is polar. In a covalent bond, unpaired valence electrons pair up in such a way that the atoms complete their outer shells.

Electron Dot Diagrams Showing Unpaired Valence Electrons (Note: When bonding occurs, molecular orbitals form. As a result, the two electrons that are normally paired in the lowest energy orbital move into separate orbitals)

\[
\begin{align*}
\text{Li} & : \\
\text{Be} & : \\
\text{B} & : \\
\text{C} & : \\
\text{N} & : \\
\text{O} & : \\
\text{F} & : \\
\text{Ne} & : \\
\end{align*}
\]

Pairing Electrons:

Nonpolar Covalent Bond: \( \text{Cl}^0 + \text{Cl}^0 \rightarrow \text{Cl}_2 \)

\[
\begin{align*}
\text{Cl} & : \\
\text{Cl} & : \\
\end{align*}
\]

Polar Covalent Bond: \( \text{H}^+ + \text{Cl}^- \rightarrow \text{HCl} \)

\[
\begin{align*}
\text{H} & : \\
\text{Cl} & : \\
\end{align*}
\]

Based on your understanding of covalent bonds, answer the questions below.

1. Draw electron dot diagrams for hydrogen and oxygen.

2. Draw electron dot diagrams showing the pairing of electrons to form water from hydrogen and oxygen. All outer shells should be complete.

3. Are the bonds in water polar or nonpolar. How do you know?
AIM: How is the shape of the molecule related to the group number?

<table>
<thead>
<tr>
<th>Elements from Groups</th>
<th>Arrangement of electrons</th>
<th>Name of molecular shape</th>
<th>Symmetrical or asymmetrical</th>
<th>Polar or nonpolar</th>
<th>Examples/general formula</th>
</tr>
</thead>
<tbody>
<tr>
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<td></td>
<td></td>
</tr>
</tbody>
</table>
Aim: How does the shape of a molecule affect its polarity?

<table>
<thead>
<tr>
<th>Symmetrical molecules</th>
<th>Asymmetrical molecules</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{BeH}_2 )</td>
<td>( \text{HBr} )</td>
</tr>
<tr>
<td>( \text{CO}_2 )</td>
<td>( \text{NH}_3 )</td>
</tr>
<tr>
<td>( \text{CH}_4 )</td>
<td>( \text{H}_2\text{O} )</td>
</tr>
<tr>
<td>( \text{BF}_3 )</td>
<td>( \text{H}_2\text{S} )</td>
</tr>
</tbody>
</table>

\[ \text{H—Be—H} \]
\[ \text{O = C = O} \]
\[ \text{H—C—H} \]
\[ \text{F} \]
\[ \text{H—Br} \]

CONCLUSION:
Using VSEPR Theory, name and sketch the shape of the following molecules.

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1.</td>
<td>N₂</td>
</tr>
<tr>
<td>2.</td>
<td>H₂O</td>
</tr>
<tr>
<td>3.</td>
<td>CO₂</td>
</tr>
<tr>
<td>4.</td>
<td>NH₃</td>
</tr>
<tr>
<td>5.</td>
<td>CH₄</td>
</tr>
<tr>
<td>6.</td>
<td>SO₃</td>
</tr>
<tr>
<td>7.</td>
<td>HF</td>
</tr>
<tr>
<td>8.</td>
<td>CH₃OH</td>
</tr>
<tr>
<td>9.</td>
<td>H₂S</td>
</tr>
<tr>
<td>10.</td>
<td>I₂</td>
</tr>
<tr>
<td>11.</td>
<td>CHCl₃</td>
</tr>
<tr>
<td>12.</td>
<td>O₂</td>
</tr>
</tbody>
</table>
## POLARITY OF MOLECULES

Determine whether the following molecules are polar or nonpolar.

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1.</td>
<td>N₂</td>
</tr>
<tr>
<td>7.</td>
<td>HF</td>
</tr>
<tr>
<td>2.</td>
<td>H₂O</td>
</tr>
<tr>
<td>8.</td>
<td>CH₃OH</td>
</tr>
<tr>
<td>3.</td>
<td>CO₂</td>
</tr>
<tr>
<td>9.</td>
<td>H₂S</td>
</tr>
<tr>
<td>4.</td>
<td>NH₃</td>
</tr>
<tr>
<td>10.</td>
<td>I₂</td>
</tr>
<tr>
<td>5.</td>
<td>CH₄</td>
</tr>
<tr>
<td>11.</td>
<td>CHCl₃</td>
</tr>
<tr>
<td>6.</td>
<td>SO₃</td>
</tr>
<tr>
<td>12.</td>
<td>O₂</td>
</tr>
</tbody>
</table>
Recognizing Polar Molecules

To determine if a compound is polar, you must consider the electronegativity difference within each bond and the three dimensional shape of the compound. If the electronegativity difference is greater than 1.7 or close to zero, the compound is not polar. Electronegativity differences above 1.7 are found in ionic compounds. Electronegativity differences around zero are found in molecules with nonpolar bonds. Electronegativity differences between 0.4 and 1.7 are found in molecules with polar bonds. These molecules can be polar or nonpolar depending on their shapes. Molecules with polar bonds distributed symmetrically are nonpolar. Asymmetrical molecules with polar bonds are polar. Water is polar. An imaginary line can be drawn through a water molecule separating the positive pole from the negative pole. This is because the charges are distributed asymmetrically. Carbon dioxide is nonpolar because the electronegative oxygens are distributed symmetrically around the carbon (O=C=O).

Determine if each of the compounds listed below, IONIC, POLAR, or NONPOLAR as follows: [1] determine the types of bonds. [2] draw electron dot diagrams to determine the shape.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Type of Bond: IONIC, POLAR, or NONPOLAR</th>
<th>Electron Dot Diagram</th>
<th>Type of Compound: IONIC, POLAR, or NONPOLAR</th>
<th>Compound</th>
<th>Type of Bond: IONIC, POLAR, or NONPOLAR</th>
<th>Electron Dot Diagram</th>
<th>Type of Compound: IONIC, POLAR, or NONPOLAR</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td></td>
<td></td>
<td></td>
<td>CCl₄</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CH₄</td>
<td></td>
<td></td>
<td></td>
<td>CH₃Cl</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Cl₂</td>
<td></td>
<td></td>
<td></td>
<td>N₂</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>KBr</td>
<td></td>
<td></td>
<td></td>
<td>H₂S</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NH₃</td>
<td></td>
<td></td>
<td></td>
<td>NaBr</td>
<td></td>
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</tr>
</tbody>
</table>

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1. A molecular compound is formed when a chemical reaction occurs between atoms of
   1) chlorine and sodium
   2) chlorine and yttrium
   3) oxygen and hydrogen
   4) oxygen and magnesium

2. What is the total number of electron pairs shared between the two atoms in an O₂ molecule?
   1) 1  2) 2  3) 6  4) 4

3. What is the total number of electrons shared in a double covalent bond?
   1) 1  2) 2  3) 3  4) 4

4. Which compound has both ionic and covalent bonding?
   1) CaCO₃  2) CH₂Cl₂
   3) CH₃OH  4) CaH₂O₆

5. The nitrogen atoms in a molecule of N₂ share a total of
   1) one pair of electrons
   2) one pair of protons
   3) three pairs of electrons
   4) three pairs of protons

6. Which formula represents a molecular compound?
   1) HI  2) KI  3) KCl  4) LiCl

7. Which formula represents a molecular compound?
   1) Kr  2) LiOH
   3) N₂O₄  4) NaI

8. Which characteristic is a property of molecular substances?
   1) good heat conductivity
   2) good electrical conductivity
   3) low melting point
   4) high melting point

9. What is the maximum number of covalent bonds that a carbon atom can form?
   1) 1  2) 2  3) 3  4) 4

10. In the diagram of an ammonium ion to the right, why is bond A considered to be a coordinate covalent bond?
    1) Hydrogen provides a pair of electrons to be shared with nitrogen.
    2) Nitrogen provides a pair of electrons to be shared with hydrogen.
    3) Hydrogen transfers a pair of electrons to the nitrogen.
    4) Nitrogen transfers a pair of electrons to hydrogen.

11. Which molecule has a nonpolar covalent bond?
    1) H-H  2) H-N-H
    3) H-O-H  4) H-Cl

12. The chemical bond between which two atoms is most polar?
    1) C-N  2) H-H  3) S-Cl  4) Si-O
13. Which formula represents a nonpolar molecule containing polar covalent bonds?
1) H₂O  2) CCl₄  3) NH₃  4) H₂

14. The bonds between hydrogen and oxygen in a water molecule are classified as
1) polar covalent  2) nonpolar covalent  3) ionic  4) metallic

15. Which type of molecule is CF₄?
1) polar, with a symmetrical distribution of charge  2) polar, with an asymmetrical distribution of charge
3) nonpolar, with a symmetrical distribution of charge  4) nonpolar, with an asymmetrical distribution of charge

16. Given the formula representing a molecule:

H – C ≡ C – H

The molecule is
1) symmetrical and polar  2) symmetrical and nonpolar  3) asymmetrical and polar  4) asymmetrical and nonpolar

17. Which formula represents a nonpolar molecule?
1) HCl  2) H₂O  3) NH₃  4) CH₄

18. Why is a molecule of CO₂ nonpolar even though the bonds between the carbon atom and the oxygen atoms are polar?
1) The shape of the CO₂ molecule is symmetrical.
2) The shape of the CO₂ molecule is asymmetrical.
3) The CO₂ molecule has a deficiency of electrons.
4) The CO₂ molecule has an excess of electrons.

19. Which formulas represent two polar molecules?
1) CO₂ and HCl  2) CO₂ and CH₄
3) H₂O and HCl  4) H₂O and CH₄

20. Which formula represents a polar molecule?
1) Br₂  2) CO₂  3) CH₄  4) NH₃
BAN DIHYDROGEN MONOXIDE!

Dihydrogen monoxide is colorless, odorless, tasteless, and kills uncounted thousands of people every year. Most of these deaths are caused by accidental inhalation of DHMO, but the dangers of dihydrogen monoxide do not end there.

Prolonged exposure to its solid form causes severe tissue damage. Symptoms of DHMO ingestion can include excessive sweating and urination, and possibly a bloated feeling, nausea, vomiting and body electrolyte imbalance. For those who have become dependent, DHMO withdrawal means certain death.

Dihydrogen monoxide:
- is also known as hydroxyl acid, and is the major component of acid rain.
- contributes to the "greenhouse effect."
- may cause severe burns.
- contributes to the erosion of our natural landscape.
- accelerates corrosion and rusting of many metals.
- may cause electrical failures and decreased effectiveness of automobile brakes.
- has been found in excised tumors of terminal cancer patients.

Contamination is reaching epidemic proportions!

Quantities of dihydrogen monoxide have been found in almost every stream, lake, and reservoir in America today. But the pollution is global, and the contaminant has even been found in Antarctic ice. DHMO has caused millions of dollars of property damage in the midwest, and recently California.

Despite the danger, dihydrogen monoxide is often used:
- as an industrial solvent and coolant.
- in nuclear power plants.
- in the production of styrofoam.
- as a fire retardant.
- in many forms of cruel animal research.
- in the distribution of pesticides.
- as an additive in certain "junk-foods" and other food products.

Even after washing, produce remains contaminated by this chemical.

Companies dump waste DHMO into rivers and the ocean, and nothing can be done to stop them because this practice is still legal. The impact on wildlife is extreme, and we cannot afford to ignore it any longer!

The American government has refused to ban the production, distribution, or use of this damaging chemical due to its "importance to the economic health of this nation." In fact, the navy and other military organizations are conducting experiments with DHMO, and designing multi-billion dollar devices to control and utilize it during warfare situations. Hundreds of military research facilities receive tons of it through a highly sophisticated underground distribution network. Many store large quantities for later use.
AIM: Writing chemical formulas for molecular compounds.

Guidelines:

Problems:

1) Nitrogen trifluoride
2) Carbon monoxide
3) Carbon tetrafluoride
4) Tetraphorous decaoxide
5) Sulfur hexafluoride
6) Dihydrogen monoxide
7) Dinitrogen tetrahydride
8) Nitrogen trihydride
9) Nitrogen monoxide
10) Dinitrogen monoxide
11) Sulfur dioxide
Directions: Write each compound in the space provided.

1. _____________________________  Sodium sulfide

2. _____________________________  Magnesium chloride

3. _____________________________  Ammonium fluoride

4. _____________________________  Sodium acetate

5. _____________________________  Silver bromide

6. _____________________________  Calcium oxide

7. _____________________________  Potassium oxide

8. _____________________________  Calcium nitrate

9. _____________________________  Lithium sulfate

10. _____________________________ Potassium phosphate

11. _____________________________ Potassium hydroxide

12. _____________________________ Magnesium chlorate
13. 

14. Carbon monoxide

15. Sulfur dioxide

16. Iron(III) nitrate

17. Carbon dioxide

18. Sulfur trioxide

19. Lead(II) oxide

20. Lead(IV) oxide

21. Sb(OH)$_3$

22. Sb(OH)$_5$

23. Co$_2$(SO$_4$)$_3$

24. CoSO$_4$
Nonmetals are two-faced elements! Although they normally have negative oxidation states, nonmetals can behave like metals and have positive oxidation states. As a result, two nonmetals can combine to form compounds. When two nonmetals combine, they form covalent bonds. The nonmetal with the lower electronegativity behaves like a metal and has a positive oxidation state. In carbon dioxide (CO₂), the carbon behaves like a metal while the oxygen behaves like a nonmetal. The metal is written first in the name and the formula. The name of the metal is the same as the name of the element (C = carbon, C = carbon). If there is more than one atom of the metal, the number of atoms is indicated with a prefix. (See the list of prefixes below.) The nonmetal is written last in the name and formula. The name of the nonmetal is the same as the name of the element minus the final syllable or two, plus IDE (O = oxygen, O²⁻ = oxide). The number of nonmetal atoms is indicated with a prefix (even when there is only one). Writing formulas for these compounds is easy, because the prefix tells the subscript.

Examples

Cl₂O₃ = dichlorine trioxide
SiF₄ = silicon tetrafluoride

Name the following binary covalent compounds.

1. BrCl₅
2. SO₃
3. P₂O₃
4. As₂P₃
5. IF₇
6. SeS₂
7. SO₂
8. CO
9. SBr₆
10. N₂O₅

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**Naming More Compounds**

Write the correct name of the compound on the space provided for each of the formulas listed below.

<table>
<thead>
<tr>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ca$_3$(PO$_4$)$_2$</td>
<td>31. KCl</td>
</tr>
<tr>
<td>PbI$_4$</td>
<td>32. MgI$_2$</td>
</tr>
<tr>
<td>Al(HSO$_4$)$_3$</td>
<td>33. NaHSO$_4$</td>
</tr>
<tr>
<td>Fe(OH)$_3$</td>
<td>34. Hg$_3$(PO$_4$)$_2$</td>
</tr>
<tr>
<td>Cl$_4$</td>
<td>35. Ni(ClO$_3$)$_2$</td>
</tr>
<tr>
<td>Cs$_2$SO$_4$</td>
<td>36. CdF$_2$</td>
</tr>
<tr>
<td>LiHCO$_3$</td>
<td>37. SnF$_2$</td>
</tr>
<tr>
<td>BaSO$_4$</td>
<td>38. CO$_2$</td>
</tr>
<tr>
<td>AlF</td>
<td>39. NaCl</td>
</tr>
<tr>
<td>NH$_4$NO$_3$</td>
<td>40. Sb(NO$_3$)$_3$</td>
</tr>
<tr>
<td>Cu$_2$O</td>
<td>41. Sn(CO$_3$)$_2$</td>
</tr>
<tr>
<td>Hg$_2$CO$_3$</td>
<td>42. KHSO$_4$</td>
</tr>
<tr>
<td>Ag$_2$C$_2$O$_4$</td>
<td>43. AsI$_3$</td>
</tr>
<tr>
<td>Cu(SCN)$_2$</td>
<td>44. NH$_3$OH</td>
</tr>
<tr>
<td>LiOH</td>
<td>45. SiCl$_4$</td>
</tr>
<tr>
<td>Ag$_2$S</td>
<td>46. NH$_4$ClO$_3$</td>
</tr>
<tr>
<td>Rb$_2$N</td>
<td>47. Cr$_2$(C$_2$O$_4$)$_3$</td>
</tr>
<tr>
<td>FeSO$_4$</td>
<td>48. NiF$_2$</td>
</tr>
<tr>
<td>ZnBr$_2$</td>
<td>49. SO$_2$</td>
</tr>
<tr>
<td>Pb(CrO$_4$)$_2$</td>
<td>50. BiCO$_3$</td>
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<tr>
<td>MnCr$_2$O$_7$</td>
<td>51. As$_2$O$_5$</td>
</tr>
<tr>
<td>SrH$_2$</td>
<td>52. CdO$_2$</td>
</tr>
<tr>
<td>Sr(CH$_3$COO)$_2$</td>
<td>53. (NH$_4$)$_2$Cr$_2$O$_7$</td>
</tr>
<tr>
<td>CS$_2$</td>
<td>54. KClO$_3$</td>
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<tr>
<td>MnO$_2$</td>
<td>55. SO$_3$</td>
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<tr>
<td>K$_2$P</td>
<td>56. Zn(NO$_3$)$_2$</td>
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<td>Na$_2$S$_2$O$_3$</td>
<td>57. C$_2$P$_4$</td>
</tr>
<tr>
<td>BaS</td>
<td>58. SnO$_2$</td>
</tr>
<tr>
<td>H$_3$PO$_3$</td>
<td>59. NH$_4$Br</td>
</tr>
<tr>
<td>FrMnO$_4$</td>
<td>60. Na$_2$O</td>
</tr>
</tbody>
</table>
When nonmetals chemically bond they do so by sharing electrons. The bond is called a covalent bond. When an active metal and a nonmetal bond, the active metal transfers one or more electrons to the nonmetal. This bond is called an ionic bond. Ionic compounds (except for bases) are also called salts.

Classify the following compounds as ionic or covalent.

1. CaCl₂
2. CO₂
3. H₂O
4. BaSO₄
5. O₂
6. NaF
7. Na₂CO₃
8. S₈
9. SO₃
10. LiBr
11. MgO
12. C₂H₅OH
13. HCl
14. N₂
15. NaOH
16. NO₂
17. Al₂O₃
18. FeCl₃
19. P₂O₅
20. N₂O₃
21. H₂
22. K₂O
23. KI
24. P₄
25. CH₄
26. NaCl

Draw an electron shell diagram of the ionic compound calcium oxide, CaO.

Draw an electron shell diagram of the covalent compound methane, CH₄.
Answer the questions below by circling the number of the correct response.

1. Barium combines by (1) gaining two electrons, (2) losing two electrons, (3) sharing two electrons, (4) sharing 3 electrons.

2. Which of the following is the correct electron dot diagram for nitrogen?

   ![N: N: N: N:](image)

   (1) (2) (3) (4)

3. In water, the bond between hydrogen and oxygen is (1) ionic, (2) polar covalent, (3) nonpolar covalent, (4) nonpolar noncovalent.

4. Which of the following occurs during covalent bonding?
   (1) Electrons are lost. (2) Electrons are gained. (3) Valence electrons fall from the excited state to the ground state. (4) Unpaired electrons form pairs.

5. Which of the following is an example of a substance with a nonpolar covalent bond? (1) HCl (2) Cl₂ (3) HClO₂ (4) NaCl

6. The electronegativity of sulfur is (1) 1.6, (2) 2.39, (3) 2.6, (4) 3.2.

7. Which of the following elements has the highest electronegativity? (1) fluorine (2) chlorine (3) barium (4) hydrogen

8. The formula for magnesium fluoride is MgF₂. The best explanation for this fact is that when they combine (1) each of two magnesium atoms lose an electron and a fluorine atom gains two, (2) a magnesium atom loses two electrons and each of two fluorine atoms gains one, (3) a magnesium atom shares two electrons with two fluorine atoms, (4) each of two magnesium atoms share an electron with a fluorine atom.

9. When calcium combines, it usually (1) loses two electrons, (2) gains six electrons, (3) shares two electrons, (4) shares six electrons.

10. What is the maximum number of atoms carbon can combine with at once? (1) 1 (2) 2 (3) 3 (4) 4

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### Drawing Lewis Diagrams for Ionic & Covalent Bonds

<table>
<thead>
<tr>
<th>Chemical formula</th>
<th>Compound name</th>
<th>Type of Bond</th>
<th>Lewis Dot Diagram</th>
</tr>
</thead>
<tbody>
<tr>
<td>F₂</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>HCl</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>K₂S</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>H₂O</td>
<td>Nitrogen</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Carbon dioxide</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Barium chloride</td>
<td></td>
<td></td>
</tr>
<tr>
<td>NH₃</td>
<td>Oxygen</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Bonding Summary
The Properties of Metals

Until about 5,000 B.C., people made tools and implements from stone. Around 5,000 B.C., it was probably noticed that melted copper ran from green ore in a pottery kiln. Around that same time, gold was discovered and used for decorations. Metals became prized for jewelry because of their luster. By 3500 B.C., humankind discovered that melting copper together with tin formed a harder metal, heralding the start of the Bronze Age. Around 1500 B.C., technology leapt forward once again when the hotter ovens of that age enabled the extraction of iron, an even harder and more abundant metal, from its ore. In many ways, we are still in the Iron Age. Modern civilization depends on alloys of iron for its bridges, skyscrapers, and automobiles. The physical properties of metals that make them so useful are due to metallic bonding which makes them both strong and flexible.

Read the description of metallic bonding below, and answer the questions that follow based on your knowledge of chemistry and metals in particular.

1. Why are metallic bonds both strong and flexible?

   ______________________________________________________

   ______________________________________________________

   ______________________________________________________

2. Why are metals able to conduct both heat and electricity well?

   ______________________________________________________

   ______________________________________________________

   ______________________________________________________

3. The valence electrons of metals jump easily to a higher energy orbital when light shines on them. Then they fall emitting the excess energy as light. Which property of metals is explained by this?

   ______________________________________________________

   ______________________________________________________
Metallic Bonds and Intermolecular Forces

**Aim**
- Examine bonds that are not chemical bonds

**Notes**

**Metallic bonds**

- **Formation**
  - in metals, electrons are easily lost or transferred
  - the electrons in metallic substances are not always associated with any particular atom
  - as a result, the particles of a metal are usually positive ions surrounded by mobile electrons to which they are attracted

- **Properties**
  - strong bonds result in high melting points
  - mobile electrons result in luster, flexibility, and good conductivity

**Intermolecular attractions** - forces of attraction between particles that are not chemically bonded

- **Dipole-dipole attraction**
  - Dipole - a polar molecule, or a molecule with an asymmetric, or unequal, distribution of charge causing one end of the molecule to be positive while the other is negative
  - Definition - force of attraction between the positive end of one dipole and the negative end of another

- **Hydrogen bonding**
  - Definition - an intermolecular force linking an electronegative hydrogen that is covalently bonded to a small electronegative element such as oxygen, nitrogen, or fluorine, to another electronegative element of the same or another molecule

- Evidence - uncharacteristically high boiling point of water

![Boiling Points of Related Compounds](image-url)
## Intermolecular Forces

<table>
<thead>
<tr>
<th>Type</th>
<th>Definition</th>
<th>Example</th>
<th>Strength</th>
</tr>
</thead>
<tbody>
<tr>
<td>Dipoles</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hydrogen Bonding</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Dispersion Forces (van der walls)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Molecule-Ion Attraction</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
The Ties that Bind

Pure substances can be held together by ionic bonds, covalent bonds, metallic bonds, or intermolecular forces. Most materials are mixtures, and are held together by a mixture of these forces. Figuring out what holds things together takes some serious analysis. If the substance is pure and you know the formula, you can figure out the electronegativity difference. If it is 1.7 or greater, than it is ionic. All ionic substances are crystalline solids. Diamonds are also crystalline solids, but they are made of pure carbon. What does that tell you about the electronegativity difference? Electronegativity differences below 1.7 are covalent. Large crystals such as diamond or sand (SiO$_2$) that have a network of covalent bonds are called macromolecules or network solids. Smaller compounds containing covalent bonds are called molecules. The molecules of a substance may be attracted to each other to form solids or liquids by intermolecular forces. These are often called molecular compounds. Molecular solids are softer than covalent solids (network solids) and ionic solids, because intermolecular forces are weaker than chemical bonds. The flow chart to the right shows one way of classifying the types of forces that hold substances together.

Once you have determined that a material is held together by intermolecular forces, this can be further refined. If the substance is polar, it is held together by dipole-dipole attractions. If the polar substance contains hydrogen atoms attached to either oxygen, nitrogen, or fluorine atoms, it forms especially strong dipole-dipole attractions called a hydrogen bonds. Hydrogen bonds are responsible for the three-dimensional shapes of many proteins because the large protein molecule folds in such a way that hydrogens in one part of the molecule are close to oxygens or nitrogens in another part of the molecule.
Below are some familiar materials. Based on the reading and your knowledge of chemistry, state whether samples of these materials are held together by ionic bonds, covalent bonds, metallic bonds, dipole-dipole attractions, hydrogen bonds, or other intermolecular forces.

1. Water [H₂O(ℓ)]
2. Table sugar [C₁₂H₂₂O₁₁(s)]
3. Table salt [NaCl(s)]
4. Iron railing [Fe(s)]
5. Liquid oxygen [O₂(g)]
6. Diamond [C(s)]
7. Salt substitute [KI(s)]
8. Alcohol [CH₃CH₂OH(ℓ)]
9. Chlorine [Cl₂(g)]
10. Gasoline [C₈H₁₈(ℓ)]
11. Gold [Au(s)]
12. Rust [Fe₂O₃(s)]
13. Tarnish [Ag₂S(s)]
14. Tooth enamel [Ca₅(PO₄)₃(s)]
15. Copper wire [Cu(s)]
Complete the observations below and answer the questions that follow.

- Fill a 50 mL beaker most of the way with water.
- Continue adding water one drop at a time
- Note how high the water can go without spilling.

1. Describe what happened when you kept adding water to an already full beaker? Draw a picture of how it looked in the space to the right. Why does this happen?

2. The graph to the right shows the boiling point of compounds of hydrogen and members of the oxygen family.
   a. What is the electronegativity difference in each compound?
   b. How can the differences in the boiling points be explained?

3. What holds many solids and liquids together?
Chemistry: Form Ls4.4A

Metallic Bonds and Intermolecular Forces

Answer the questions below by circling the number of the correct response

1. Which substance will conduct electricity in both the solid phase and the liquid phase?
   (1) AgCl   (2) H₂   (3) Ag   (4) HCl

2. Hydrogen bonds are strongest between molecules of
   (1) HBr(g)   (2) H₂(g)   (3) HF(g)   (4) HCl(g)

3. Which molecule is a dipole?
   (1) H₂   (2) N₂   (3) CH₄   (4) HCl

4. The strongest hydrogen bonds are formed between molecules of
   (1) H₂Te   (2) H₂Se   (3) H₂O   (4) H₂S

5. What type of bonds are present in a strip of magnesium ribbon?
   1 covalent   2 ionic   3 metallic   4 van der Waals

6. Hydrogen bonds are most likely to exist between molecules of
   (1) H₂   (2) CH₄   (3) HF   (4) H₂O

7. Which substance, in the solid state, is the best conductor of electricity?
   (1) Ag   (2) I₂   (3) NaCl   (4) CO₂

8. Which is the predominate type of attraction between molecules of HF in the liquid state?
   1 hydrogen bonding   2 ionic bonding
   3 covalent bonding   4 dipole bonding

9. Which substance exists as a metallic crystals
   (1) Ar   (2) Au   (3) SiO₂   (4) CO₂

10. Mobile electrons are a distinguishing characteristic of
    1 an ionic bond   2 a metallic bond
    3 an electrovalent bond   4 a covalent bond

11. Which kinds of bonds are found in a sample of H₂O(s)?
    1 hydrogen bonds, only
    2 covalent bonds, only
    3 both ionic and hydrogen bonds
    4 both covalent and hydrogen bonds

12. Which substance is made up of molecules that are dipoles?
    (1) N₂   (2) H₂O   (3) CH₄   (4) CO₂

13. Which element consists of positive ions immersed in a "sea" of mobile electrons?
    1 sulfur   2 nitrogen
    3 calcium   4 chlorine

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Naming: Putting it all Together

Below is a flow chart showing the steps to follow when naming a compound from the formula. Following the flow chart will enable you to determine when to use the stock system, when to use a set of prefixes, and when to look up the names of polyatomic ions in Table E.

The compounds below are of several different types. Use the flow chart to determine the naming system to use and name each compound shown below.

1. Fe(NO₃)₃
2. Na₃S₂O₃
3. P₂O₅
4. BaBr₂
5. Mn₅(Cr₂O₇)²⁻
6. CaCl₂
7. (NH₄)₂S
8. CuF
9. Br₂O
10. HgSO₄
11. Al₂O₃
12. SCl₆
13. IF₇
14. Cr(CO₃)₃
15. KNO₂
16. AuP

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Is defined as the force of attraction for electrons between nuclei.

It can be:

**Metallic**
- Which is a "sea of mobile electrons".
  - Ex: Ca (s) and Au (s)
    - A Metal

**Ionic**
- Which is a transfer of electrons.
  - Ex: NaCl and KF
    - A Metal & Nonmetal

**Covalent**
- Which is a sharing of electrons.
  - A special kind is:
    - Coordinate Covalent
      - Which is a covalent bond in which one atom donates both electrons.
        - Ex: \( \text{NH}_4^+ \)
          - Formed in polyatomic ions.
    - Nonpolar
      - Which is an equal sharing of electrons
        - Ex: H₂, N₂
          - Two same nonmetals
    - Polar
      - Which is an unequal sharing of electrons
        - Ex: H₂O, NH₃
          - Two different nonmetals
1. Which formula represents a molecule having a nonpolar covalent bond?
   A) \( \text{H} \)  
   B) \( \text{H} \)  
   C) \( \text{H} \)  
   D) \( \text{H} \)

2. The chemical bond between which two atoms is most polar?
   A) C-N  
   B) H-H  
   C) S-Cl  
   D) Si-O

3. Which formula represents a nonpolar molecule containing polar covalent bonds?
   A) \( \text{H}_2\text{O} \)  
   B) \( \text{CCl}_4 \)  
   C) \( \text{NH}_3 \)  
   D) \( \text{H}_2 \)

4. The bonds between hydrogen and oxygen in a water molecule are classified as
   A) polar covalent  
   B) nonpolar covalent  
   C) ionic  
   D) metallic

5. Which type of molecule is \( \text{CF}_4 \)?
   A) polar, with a symmetrical distribution of charge  
   B) polar, with an asymmetrical distribution of charge  
   C) nonpolar, with a symmetrical distribution of charge  
   D) nonpolar, with an asymmetrical distribution of charge

6. Which compound has molecules that form the strongest hydrogen bonds?
   A) HI  
   B) HBr  
   C) HF  
   D) HCl

7. Which electron-dot diagram represents a molecule that has a polar covalent bond?
   A) \( \text{H:Cl:} \)  
   B) \( \text{Li}^+ \left[ \text{Cl}_2^- \right] \)  
   C) \( \text{Cl}_2 \times \text{Cl}_2 \times \)  
   D) \( \text{K}^+ \left[ \text{Cl}_2^- \right] \)

8. Which molecule contains a polar covalent bond?
   A) \( \text{H:Cl:} \)  
   B) \( \text{H:O:} \)  
   C) \( \text{H:N:H} \)  
   D) \( \text{N=N:} \)

9. Given the formula representing a molecule:
   \[ \text{H-C≡C-H} \]
   The molecule is
   A) symmetrical and polar  
   B) symmetrical and nonpolar  
   C) asymmetrical and polar  
   D) asymmetrical and nonpolar

10. Which electron-dot structure represents a non-polar molecule?
    A) \( \text{H:Cl:} \)  
    B) \( \text{H:O:} \)  
    C) \( \text{H:N:H} \)  
    D) \( \text{H:O:} \)

11. Two fluorine atoms are held together by a covalent bond. Which statement correctly describes this bond?
    A) It is polar and forms a polar molecule.  
    B) It is polar and forms a nonpolar molecule.  
    C) It is nonpolar and forms a polar molecule.  
    D) It is nonpolar and forms a nonpolar molecule.

12. Given a formula for oxygen:
    \[ \text{O=O} \]
    What is the total number of electrons shared between the atoms represented in this formula?
    A) 1  
    B) 2  
    C) 8  
    D) 4
13. What is the total number of pairs of electrons shared in a molecule of N\textsubscript{2}?

A) one pair  
B) two pairs  
C) three pairs  
D) four pairs

14. What is the total number of electrons shared in the bonds between the two carbon atoms in the molecule shown below?

\[ \text{H} - \text{C} = \text{C} - \text{H} \]

A) 6  
B) 2  
C) 3  
D) 8

15. Which compound contains only covalent bonds?

A) NaOH  
B) Ba(OH)\textsubscript{2}  
C) Ca(OH)\textsubscript{2}  
D) CH\textsubscript{3}OH

16. Which pair of atoms is held together by a covalent bond?

A) HCl  
B) LiCl  
C) NaCl  
D) KCl

17. Which compound contains both ionic and covalent bonds?

A) HCl(g)  
B) NaCl(s)  
C) NH\textsubscript{4}Cl(s)  
D) CCl\textsubscript{4}(l)

18. Which is the correct electron-dot formula for a hydrogen molecule at STP?

A) H⁻  
B) H⁺  
C) H·H  
D) H·H

19. Which characteristic is a property of molecular substances?

A) good heat conductivity  
B) good electrical conductivity  
C) low melting point  
D) high melting point

20. In the diagram of an ammonium ion to the right, why is bond \textit{A} considered to be a coordinate covalent bond?

\[ \text{H}^+ \overset{\equiv}{\text{N}}\overset{\equiv}{\text{H}} \]

A) Hydrogen provides a pair of electrons to be shared with nitrogen.  
B) Nitrogen provides a pair of electrons to be shared with hydrogen.  
C) Hydrogen transfers a pair of electrons to the nitrogen.  
D) Nitrogen transfers a pair of electrons to hydrogen.
Base your answer to questions 1 through 3 on the table below.

### Physical Properties of Four Gases

<table>
<thead>
<tr>
<th>Name of Gas</th>
<th>Molecular Structure</th>
<th>Boiling Point (K) at 1 Atm</th>
<th>Density (g/L) at STP</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen</td>
<td>H–H</td>
<td>20</td>
<td>0.0899</td>
</tr>
<tr>
<td>hydrogen chloride</td>
<td>H–Cl</td>
<td>188</td>
<td>1.64</td>
</tr>
<tr>
<td>hydrogen bromide</td>
<td>H–Br</td>
<td>207</td>
<td>?</td>
</tr>
<tr>
<td>hydrogen iodide</td>
<td>H–I</td>
<td>237</td>
<td>5.86</td>
</tr>
</tbody>
</table>

1. Explain, in terms of molecular polarity, why hydrogen chloride is more soluble than hydrogen in water under the same conditions of temperature and pressure.

2. Explain, in terms of intermolecular forces, why hydrogen has a lower boiling point than hydrogen bromide.

3. Explain, in terms of electronegativity difference, why the bond in H–Cl is more polar than the bond in H–I.

4. Draw a Lewis electron-dot diagram for a molecule of phosphorus trichloride, PCl₃

Base your answer to questions 5 and 6 on the balanced equation below.

\[ 2\text{Na(s)} + \text{Cl}_2 \rightarrow 2\text{NaCl(s)} \]

5. Draw a Lewis electron-dot diagram for a molecule of chlorine, Cl₂.

6. Explain, in terms of electrons, why the bonding in NaCl is ionic.
7. Base your answer to the following question on the information below.

Each molecule listed below is formed by sharing electrons between atoms when the atoms within the molecule are bonded together.

Molecule A: Cl₂
Molecule B: CCl₄
Molecule C: NH₃

Explain why CCl₄ is classified as a nonpolar molecule.

8. Bromine is the only liquid nonmetallic element at room temperature. It is a heavy, mobile, reddish-brown liquid, volatilizing readily at room temperature to a red vapor with a strong disagreeable odor, resembling chlorine, and having a very irritating effect on the eyes and throat; it is readily soluble in water or carbon disulfide, forming a red solution, is less active than chlorine but more so than iodine; it unites readily with many elements and has a bleaching action; when spilled on the skin it produces painful sores. It presents a serious health hazard, and maximum safety precautions should be taken when handling it.

   a) Draw the electron-dot diagram of a molecule of bromine, Br₂.
   b) Why does bromine have properties resembling chlorine?

9. a) Draw the structural formula for H₂O.
   b) Is this molecule polar or nonpolar? Explain your answer.

10. The natural gas delivered to consumers in the U. S. is about 95% methane (CH₄, molecular weight 16 g/mol) with the remainder being mostly ethane, propane, and carbon dioxide.

    Draw the Lewis electron-dot structure for a molecule of methane.
Bonding

What will students know and be able to do by the end of this instructional unit?

1. Define and recognize:
   a. Asymmetry
   b. Chemical Bond
   c. Covalent Bond
   d. Intermolecular Forces of Attraction
   e. Ionic Bond
   f. Metallic Bond
   g. Molecule
   h. Nonpolar
   i. Polar
   j. Polyatomic Ion
   k. Symmetry

2. Compare Ionic & molecular compounds

3. Explain properties in terms of their corresponding bond types

4. Interpret and draw lewis dot diagrams

5. Assess polarity of a bond by using their elements electronegativity values

6. Use the type of element to predict the polarity of the bond

7. Differentiate between intermolecular forces

8. Identify and differentiate polar and nonpolar molecules

9. Apply “like-dissolves-like” to real world applications

10. Identify the effects of intermolecular forces on physical properties and behavior

11. In terms of intermolecular forces, explain vapor pressure, evaporation rate and phase changes

12. Interpret VP curves in terms of boiling points, evaporation, intermolecular forces, and relative vapor pressure

13. Explain colligative properties of solutes on water’s boiling point, freezing point and vapor pressure

Key Subject Competencies

- Explain energy changes associated with forming chemical bonds.
- Describe the octet rule and how it relates to chemical bonding.
- Describe what a noble gas configuration is and determine the n.g.c. for different elements.
- Explain how ionic bonds form and the properties caused by them (ionic substances).
- Explain how covalent bonds form and the properties caused by them (network-covalent, molecular-covalent substances).
- Use the electronegativity difference between 2 elements to determine a compound’s bond type.
- Explain how metallic bonds form and the properties caused by them.
- Use element types (metal + nonmetal; only nonmetals; only metals) to determine what bond type is present in a substance.
- Describe what polyatomic ions are and how to draw them in compounds.
- Draw bond types (ionic or covalent) using Lewis Structures.
- Explain how a substance (metallic or aqueous/melted ionic) is able to conduct electricity. Explain why a substance (solid ionic or molecular or network) is NOT able to conduct electricity.
Bonding

- Explain and draw the difference between a single-covalent, double-covalent, and a triple-covalent bond.
- Determine and describe what it means for a bond to be a polar bond or a nonpolar bond, using charge symmetry.
- Determine and describe what it means for a molecule to be polar molecule or a nonpolar molecule, using charge symmetry.
- Describe what an intermolecular force of attraction is and how they affect properties such as solubility, melting point temperature, and boiling point temperature.
- Prioritize the types of intermolecular forces of attraction based on their strengths of attraction between molecules

Vocabulary
- Anion
- Ion
- Octet Rule
- Asymmetrical Molecule
- Ionic Bond
- Polar Covalent Bond
- Cation
- Ion-Molecule
- Symmetrical Molecule
- Chemical Bond
- Lewis Dot Diagram
- Triple Covalent Bond
- Covalent Bond
- Metallic Bond
- Van der Waal’s Forces
- Crystal Lattice Molecule
- Dipole-Dipole
- Multiple Covalent Bond
- Hydrogen Bond
- Nonpolar Covalent Bond
- Intermolecular Forces
- Octet
Electronegativity is a scale used to determine an atom’s attraction for an electron in the bonding process. Differences in electronegativities are used to predict whether the bond is pure covalent, polar covalent, or ionic. Molecules in which the electronegativity difference is zero are considered to be pure covalent. Those molecules that exhibit an electronegativity difference of more than zero but less than 1.7 are classified as polar covalent. Ionic crystals exist in those systems that have an electronegativity difference of more than 1.7.

The structures used to show the bonding in covalent molecules are called Lewis structures. When bonding, atoms tend to achieve a noble gas configuration. By sharing electrons, individual atoms can complete the outer energy level. In a covalent bond, an octet of electrons is formed around each atom (except hydrogen.)

To study covalent molecules, chemists find the use of models helpful. Colored wooden or plastic balls are used to represent atoms. These balls have holes drilled in them according to the number of covalent bonds they will form. The holes are bored at angles that approximate the accepted bond angles.

Sticks and springs are used to represent bonds. Single bonds are shown with sticks, while double and triple bonds are shown with two springs and three springs, respectively. While the sizes of the atoms are not proportionately correct, the models are useful to represent the arrangement of the atoms according to their bond angles.

Problem
How can we determine the type of bonds in a compound and draw and construct models of molecules?

Objectives
- Construct models to show the shapes of some covalent compounds.
- Draw a Lewis representation of the structure of some molecules.
- Compare models and Lewis structures of molecules.

Materials
- wooden or plastic molecular model set (ball and stick)
- pliers
- electronegativity tables

Safety Precautions
Always wear safety goggles and a lab apron.
Pre-Lab

1. Define covalent bond.
2. Give the electron configuration of oxygen, hydrogen, nitrogen, and carbon.
3. How many covalent bonds will each of oxygen, hydrogen, nitrogen, and carbon form?
4. Describe how electronegativity differences are used to predict whether a bond is pure covalent, polar covalent, or ionic.
5. Read the entire laboratory activity. Form a hypothesis about how to show sharing of electrons in a covalent bond in an illustration and in a model and how the type of bond is determined. Record your hypothesis on page 71.

Procedure

Part A

1. Look at your ball-and-stick model sets. Identify the different pieces that represent atoms, single bonds, double bonds, and triple bonds.
2. Select one of every different color of ball. Each hole that has been bored into the sphere represents a single chemical bond. Count the number of holes present in the different colored balls. Record your observations in Data Table 1.

Part B

1. Use an electronegativity table (see page 169 in your textbook) to determine the electronegativity difference between the two elements in the compounds in Data Table 2.
2. Use the tables on the right to determine the percentage of ionic character and bond type of each of the compounds. Record your answers on Data Table 2.

Part C

1. Construct a model for H₂.
2. Compute the electronegativity difference for the atoms in the molecule and identify the type of bond. Record your answer on Data Table 3.
3. Draw the Lewis structure for the molecule in the space provided on Data Table 3.

Table 1

<table>
<thead>
<tr>
<th>Electronegativity difference</th>
<th>Bond type</th>
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<tr>
<td>0</td>
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<td>Greater than 1.7</td>
<td>ionic</td>
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</table>

Table 2

<table>
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<tr>
<th>Electronegativity difference</th>
<th>Type of bond</th>
<th>Percent ionic character</th>
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<tbody>
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<td>0</td>
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<td>2.6</td>
<td>ionic</td>
<td>82</td>
</tr>
<tr>
<td>2.8</td>
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<td>86</td>
</tr>
<tr>
<td>3.0</td>
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<td>89</td>
</tr>
<tr>
<td>3.2</td>
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<td>92</td>
</tr>
</tbody>
</table>
Hypothesis

Cleanup and Disposal

1. Be sure all sticks and springs have been removed from the spheres.
2. Nearly reassemble the model kit.

Data and Observations

<table>
<thead>
<tr>
<th>Ball color</th>
<th>Number of holes</th>
<th>identity of element</th>
</tr>
</thead>
<tbody>
<tr>
<td>Red</td>
<td></td>
<td>oxygen</td>
</tr>
<tr>
<td>Orange</td>
<td></td>
<td>bromine</td>
</tr>
<tr>
<td>Yellow</td>
<td></td>
<td>hydrogen</td>
</tr>
<tr>
<td>Green</td>
<td></td>
<td>chlorine</td>
</tr>
<tr>
<td>Blue</td>
<td></td>
<td>nitrogen</td>
</tr>
<tr>
<td>Purple</td>
<td></td>
<td>iodine</td>
</tr>
<tr>
<td>Black</td>
<td></td>
<td>carbon</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Formula</th>
<th>Electronegativity difference</th>
<th>Percent ionic character</th>
<th>Type of bond</th>
</tr>
</thead>
<tbody>
<tr>
<td>KCl</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>K_2O</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Br_2</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>MgI_2</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>HBr</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CaCl_2</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaBr</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>MgS</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Al_2S_3</td>
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<td></td>
<td></td>
</tr>
<tr>
<td>NaCl</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>F_2</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>SO_3</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>HCl</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CO</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Analyze and Conclude

1. Observing and Inferring  Both water and carbon dioxide are triatomic molecules. Explain the meaning of triatomic.

2. Collecting and Interpreting Data  Compare the appearance of the Lewis structure for a compound with a ball and stick model of the compound.

3. Predicting  Predict the shape and Lewis structure for CBr₄.

4. Drawing a Conclusion  Explain why a formula without electronegativity data or a Lewis structure cannot be used to predict bond type.

5. Error Analysis  Compare the ball and stick models you constructed with your Lewis structures. Do any of them differ in the number of bonds? What could be some causes for the errors?

1. Explain why water is a liquid at room temperature and carbon dioxide is a gas.

2. Naphthalene (C₁₀H₈), a common ingredient in moth balls, melts at 80.2°C. Sodium chloride (NaCl), common table salt, melts at 800.7°C. What do these melting points indicate about the bonding pattern of each compound?
Three-Dimensional Models of Covalent Molecules

Pre-Lab Discussion

A single covalent bond is formed when two atoms share a pair of electrons. Each atom provides one of the electrons of the pair. If the two atoms are alike, the bond is said to be nonpolar covalent. If the atoms are unlike, one exerts a greater attractive force on the electrons, and the bond is polar covalent. More than one pair of electrons can be shared. This results in a double or triple bond.

A group of atoms held together by covalent bonds is called a molecule. Molecules can be either polar or nonpolar. If bonds are nonpolar, the molecule is nonpolar. If bonds are polar, molecules can still be nonpolar if the charge distribution throughout the molecule is symmetrical. A molecule's symmetry depends on its shape, that is, the positions in space of the atoms making up the molecule. Some possible shapes are linear, angular (bent), pyramidal, and tetrahedral. Although we represent molecules on paper as being two-dimensional for convenience, they are actually three-dimensional. By building molecular models, chemists come to understand the bonding, shapes, and polarity of even the most complex molecules.

Purpose

Build three-dimensional models of some simple covalent molecules. Predict their shapes and polarities from knowledge of bonds and molecule polarity rules.

Equipment

molecular model building set

Safety

All general lab safety rules should be followed. Always wear safety goggles and a lab apron or coat when working in the lab.

Procedure

1. Obtain a molecular model building set. Study the color code identifying the different kinds of atoms.
2. Observe that the following atoms have one hole (bonding site): hydrogen, fluorine, chlorine, bromine, and iodine. The atoms with two holes are oxygen and sulfur. A nitrogen atom has three holes, and a carbon atom has four holes.
3. Construct models of the following molecules:

- \( H_2 \) (hydrogen)
- \( HF \) (hydrogen fluoride)
- \( CH_3OH \) (methanol)
- \( CH_2O \) (carbon monoxide)
- \( CH_2Cl_2 \) (dichloromethane)
- \( N_2 \) (nitrogen)
- \( O_2 \) (oxygen)
- \( H_2S \) (hydrogen sulfide)

4. Record your observations below.

### Observations and Data

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Structural representation</th>
<th>Shape</th>
<th>Molecule polarity</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen</td>
<td>( H_2 )</td>
<td>( H - H )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>water</td>
<td>( H_2O )</td>
<td>( H - O - H )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>methane</td>
<td>( CH_4 )</td>
<td>( H - C - H )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>chlorine</td>
<td>( Cl_2 )</td>
<td>( Cl - Cl )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>ammonia</td>
<td>( NH_3 )</td>
<td>( H - N - H )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>hydrogen fluoride</td>
<td>( HF )</td>
<td>( H - F )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>ethyne</td>
<td>( C_2H_2 )</td>
<td>( H - C = C - H )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>dichloromethane</td>
<td>( CH_2Cl_2 )</td>
<td>( H - C - Cl )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>nitrogen</td>
<td>( N_2 )</td>
<td>( N = N )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>carbon dioxide</td>
<td>( CO_2 )</td>
<td>( O = C = O )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>methanol</td>
<td>( CH_3OH )</td>
<td>( H - C - O - H )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>hydrogen peroxide</td>
<td>( H_2O_2 )</td>
<td>( O - O )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>oxygen</td>
<td>( O_2 )</td>
<td>( O - O )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>hydrogen sulfide</td>
<td>( H_2S )</td>
<td>( S - H - S )</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Conclusions and Questions

1. Which molecules were nonpolar because all bonds were nonpolar?

2. Which molecules had polar covalent bonds but were nonpolar because of symmetry?

3. Which two shapes appeared to produce polar molecules?

4. Name two types of substances that do not contain molecules with covalent bonds.