Ionic and Covalent Compounds: Structures and Properties

Chemical bond: Attractive force between 2 atoms in a compound

Lewis Dot Structure: Specifies an element and uses dots to show only the valence electrons

Examples: Mg: Na

During chemical reactions, electrons can be transferred to or from atoms in order to fulfill the octet rule

Example: Na $\rightarrow$ Na$^{+1} + e^-$
Cl + e$^-$ $\rightarrow$ Cl$^{-1}$

Ionic Bond: Attractive force between oppositely charged particles (electrostatic force); results from transfer of electrons between atoms

Na$^{+} + $ Cl$^{-} \rightarrow $ NaCl

Covalent Bond: A sharing of electrons between 2 atoms to fulfill the octet rule

Examples: H + H $\rightarrow$ H$_2$; HF, H$_2$O

We can also use a dash (-) to represent a covalent bond between atoms: H-F, H-O-H

Double Bond: Bond resulting from the sharing of 2 electrons; stronger than a single bond

Example: CO$_2$ O=C=O

Triple Bond: Bond resulting from the sharing of 3 electrons; stronger than a double bond

Polarity of Covalent Bonds

Not all atoms attract electrons to the same degree

Electronegativity (E$_n$): Tendency of an atom to attract shared electrons of a covalent bond

Electronegativity increases across Periodic Table, decreases down Periodic Table

Electronegativity Values
Nonpolar Covalent Bond: Covalent bond in which the bonding electrons are shared equally by both atoms

Example: Cl-Cl, H-H

Polar Covalent Bond: Covalent bond in which the bonded electrons are shared unequally

Example: H-Cl, H-O-H

This polarization (unequal distribution of electrons) results in "partial charges" on the atoms within the covalently bonded molecule, while the net charge on the molecule remains zero.

The greater the difference in electronegativity between bonded atoms, the more polar the bond

\[ \Delta E_n = 0 \] nonpolar covalent bond
\[ \Delta E_n < 1.9 \] polar covalent bond
\[ \Delta E_n > 1.9 \] ionic bond

Writing the Names and Formulas of Ionic Compounds

Things to Know About Ionic Compounds:

• Metals tend to lose electrons during ionic bond formation
• Nonmetals tend to gain electrons during ionic bond formation
• No more than 3 electrons are transferred

• Metals lose same number of electrons as their group number
• Nonmetals gain the number of electrons equal to their group number subtracted from 8 (group number - 8)

Examples:

Li (metal, group IA)

\[ \text{Li} \rightarrow \text{Li}^+ + 1 \text{e}^- \]
Cl (nonmetal, group VIIA)

\[ \text{Cl} + 1 \text{e}^- \rightarrow \text{Cl}^- \]

The more electronegative atom acquires a partial negative charge (\(\delta^-\)); the less electronegative atom acquires a partial positive charge (\(\delta^+\))

Example: H-Cl

H has less electronegativity than Cl

Therefore electrons are more attracted to Cl atom

\(\delta^- \diamond \delta^+\)
H-Cl

Practice: I-Cl

The ratio of positive to negative ions is determined by the charges on the ions (the number of electrons transferred)

• The total positive and total negative charges in the final formulas must add up to zero

Example: Na and Cl

\[ \text{Na} \rightarrow \text{Na}^+ + \text{e}^- \]
\[ \text{Cl} + 1 \text{e}^- \rightarrow \text{Cl}^- \]

These ions will combine in a 1:1 ratio, to give an ionic compound with a net charge of zero:

NaCl
Example: Mg and F

\[ \text{Mg} \rightarrow \text{Mg}^{2+} + 2 \text{e}^- \]
\[ \text{F} + 1 \text{e}^- \rightarrow \text{F}^- \]
\[ \text{Mg}^{2+} + \text{F}^- \rightarrow \text{MgF}_2 \]

Practice: K and S

\[ \text{K} \rightarrow \text{K}^+ + 1 \text{e}^- \]
\[ \text{S} + 2 \text{e}^- \rightarrow \text{S}^2- \]

What is the formula of the ionic compound formed by these 2 ions?

\[ \text{K}_2\text{S} \]

Practice: Ca and Br

\[ \text{Ca} \rightarrow \text{Ca}^{2+} + 2 \text{e}^- \]
\[ \text{Br} + 1 \text{e}^- \rightarrow \text{Br}^- \]

What is the formula of the ionic compound formed by these 2 ions?

\[ \text{CaBr}_2 \]

\[ \text{Fe}^{3+} + \text{O}^2- \rightarrow \text{Fe}_2\text{O}_3 \]

Ionic compounds exist as crystal structures

Crystal lattice: rigid 3-dimensional arrangement of particles

Naming Ionic Compounds

Name of metallic element comes first, followed by stem of nonmetal with the suffix “ide” appended

compound name = metal + nonmetal stem + ide

Examples:

- KCl = Potassium chloride
- SrO = Strontium oxide
- Ca\(_3\)N\(_2\) = Calcium nitride

Names of Single Ions

Names for individual ions follow the same system (metal name is un-changed, non-metal suffix changes to “ide”)

Example: K\(^+\) = Potassium ion
Cl\(^-\) = Chloride ion

Some atoms (such as the transition metals) can form more than one type of charged ion:

- Cu \rightarrow Cu\(^+\) and Cu\(^2+\)
- Fe \rightarrow Fe\(^{2+}\) and Fe\(^{3+}\)
**Stock System:** Uses Roman numeral in name of chemical formula to indicate the charge on metal atom

**Example:** CuCl = Copper(I) chloride
   CuCl₂ = Copper(II) chloride

Older system uses suffixes “-ous” and “-ic” attached to root of metal name (uses non-English stem for elements with non-English names)

-ous = ion of lower charge
-ic = ion of higher charge

**Example:**
- Copper (Latin name = cuprum)
  - CuCl = Cuprous chloride
  - CuCl₂ = Cupric chloride
- Iron (Latin name = ferrum)
  - FeCl₂ = Ferrous chloride
  - FeCl₃ = Ferric chloride

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**Naming Transition Metal Ions**

(could have more than one charge)

**Naming Covalent Compounds**

Covalent bonds usually occur between non-metals

- Name of the less electronegative element first
- Stem of the name of the more electronegative element plus the “ide” suffix
- Indicate the number of each type of atom by using the Greek prefix (mono, di, tri, etc.) (omit “mono” from the first element)

**Examples:**
- N₂O₅ = Dinitrogen pentoxide
- CO₂ = Carbon dioxide, CCl₄ = Carbon tetrachloride

**Practice:** BF₃ = ?  SO₂ = ?

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**Drawing Lewis Dot Structures**

(for individual atoms or ions, polyatomic ions or covalent molecules)

**Monoatomic ion:** Consists of a single atom

**Polyatomic ion:** Group of atoms bonded together covalently that acts as a unit, with an overall charge, and forms ionic bonds with other ions

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**Names of Common Polyatomic Ions:**
Memorize those in blue, plus nitrite and sulfite
1. Establish a skeletal structure
   a. least electronegative atom in center
   b. H, F often at ends, seldom in center
   c. carbon likes to form C-C chains

2. Determine the total number of valence electrons in the entire compound
   a. add up valence electrons for each atom
   b. add an electron for each negative charge
   c. subtract an electron for each positive charge

3. Place a pair of electrons between bonded atoms

4. Move electron pairs around as needed to satisfy the octet rule for the central atom (double or triple bonds may be necessary)

5. Double check that each atom has a full octet and that the total number of electrons shown matches that calculated in step #2

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**Molecular Geometry: VSEPR Theory**

Covalent compounds exist as 3-dimensional structures; we can predict these shapes based on electron arrangement.

**VSEPR Theory:** Valence-Shell Electron Pair Repulsion theory; electron pairs in the valence shell repel each other and try to get as far away from each other as possible.

**Central Atom:** any atom in a molecule or ion that is bonded to 2 or more other atoms

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**Pick Your Poison**

![Molecules](image)

Electron pairs around a central atom orient themselves as far away from each other as possible.

**Linear:** 2 pairs of electrons around central atom, orient on opposite sides of the central atom (180° apart)

**Trigonal Planar:** 3 pairs of electrons around central atom, orient in a triangle around the central atom (120° apart)

**Tetrahedral:** 4 bonded pairs of electrons orient around the central atom in the middle (109.5° apart)

**Trigonal pyramidal:** 3 bonded pairs and 1 lone pair of electrons orient in a tetrahedron around the central atom (107° apart)

**Bent/Angular:** 2 bonded pairs of electrons and 2 lone pairs of electrons orient in a tetrahedron around the central atom (104.5° apart)
Examples of Covalent Molecular Shapes

Linear

[Diagram]

Trigonal planar

[Diagram]

Tetrahedral

[Diagram]

Trigonal Pyramid

[Diagram]

Bent/Angular

[Diagram]

When counting the number of electron pairs surrounding a central atom, remember:

- All valence shell electron pairs count, whether they are bonding pairs or lone pairs
- Double or triple bonds are counted like a single pair of electrons when predicting shapes
Determining Molecular Shape from Lewis Dot Structure

**Lewis Dot Structures and Polarity of Molecules**

Whole molecules can also be described as either polar or nonpolar

**Polar molecules:** molecules with *uneven* charge distribution; a positive end and a negative end to the molecule (dipole)

**Nonpolar molecules:** molecules whose charge distribution is symmetrical within the molecule; no positive or negative part

**Examples:**
- H₂ is a nonpolar molecule
- HCl is a polar molecule
- H₂O is a very polar molecule

Polarity affects many properties (solubility, melting point, boiling point)

**Intermolecular forces:** interactions between different molecules

**Intramolecular forces:** interactions within a molecule; the bonds between atoms that hold a molecule together

Molecules with more intermolecular forces have higher melting and boiling points (ionic, polar covalent, nonpolar covalent)

See fig. 3.6 page 102

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**Table 4.3 Molecular Structures**

<table>
<thead>
<tr>
<th>Bonded</th>
<th>Bonding Electrons/Pairs</th>
<th>Bond Angle</th>
<th>Molecular Structure</th>
<th>Example</th>
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<td>&lt;90°</td>
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