Chapter 4. Chemical Bonding: The Ionic Bond Model

Introduction to Inorganic Chemistry
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Online Tests on Following days
March 24, 2017: Test 1 (Chapters 1-3)
April 7, 2017: Test 2 (Chapters 4-5)
April 28, 2017: Test 3 (Chapters 6, 7 & 8)
May 12, 2017: Test 4 (Chapters 9, 10 & 11)
May 15, 2017: Make Up Exam: Chapters 1 - 11

Section 4.1
Chemical Bonds

A Chemical Bond
• Attractive force that holds two atoms together in a more complex unit.
• Form as a result of interactions between electrons found in the combining atoms.

Ionic Bond
• Chemical bond formed through the transfer of one or more electrons from one (metal) atom or group of atoms to another (non-metal) atom or group of atoms.
• Ionic Compound
  – A compound in which ionic bonds are present due to charged attractions between cations and anions.

Two Types of Chemical Bonds
• Ionic Bonds (metal + non-metal) Chapter 4
• Covalent Bonds (non-metal + non-metal) Chapter 5
• Metallic Bonds (metal + metal) (not discussed)

Section 4.1
Chemical Bonds

Covalent Bond
• Chemical bond formed through the sharing of one or more pairs of electrons between two non-metal atoms.
• Molecular Compound (Covalent Compound)
  – A compound in which atoms are joined through covalent bonds.
**Chemical Bonds**

**Section 4.1**

**Metallic Bond**
- Chemical bond formed through the sharing of one or more pairs of electrons between all atoms in a solid.
- **Metals:** Metallic elements
  - Metallic properties are due to metallic bonding
- **Alloys** (Metallic compounds)
  - A compound in which atoms are joined through metallic bonds.

**Bonding**
- Most bonds are not 100% ionic or 100% covalent.
- Most bonds have some degree of both ionic and covalent character.

**Section 4.2**

**Valence Electrons and Lewis Symbols**

**Valence Electron**
- An electron in the outermost electron shell of a representative element or noble-gas element.
- In these representative elements or noble gases the valence electrons are found in either s or p subshells.

**Lewis Symbol**
- Chemical symbol of an element surrounded by dots equal in number to the number of valence electrons present in atoms of the element.
Section 4.2
Valence Electrons and Lewis Symbols

Concept Check

Determine the number of valence electrons in each of the following elements:

case
selenium

1 valence electron (4s^2)

Carbon
sage

$\text{O} \quad \text{P} \quad \text{F}$

Three Important Generalizations About Valence Electrons

1. Representative elements in the same group have the same number of valence electrons.
2. The number of valence electrons for representative elements is the same as the Roman numeral periodic-table group number.
3. The maximum number of valence electrons for any element is eight.

Concept Check

Write Lewis symbols for the following elements:

O
P
F

The Octet Rule

- Certain arrangements of valence electrons are more stable than others.
- The valence electron configurations of the noble gases are considered the most stable of all valence electron configurations.
Octet Rule (G.N. Lewis)

- In forming compounds, atoms of elements lose, gain, or share electrons in such a way as to produce a noble-gas electron configuration for each of the atoms involved.

Ion

- An atom (or group of atoms) that is electrically charged as a result of the loss or gain of electrons.
  - If an atom gains one or more electrons, it becomes a negatively charged ion (anion).
  - If an atom loses one or more electrons, it becomes a positively charged ion (cation).

 Isoelectronic to Ne

Isoelectronic to Ar

Concept Check

Give the chemical symbol for each of the following ions.

a) The ion formed when a potassium atom loses one electron.  
   \[ \text{K}^+ \]

b) The ion formed when a sulfur atom gains two electrons.  
   \[ \text{S}^{2-} \]

Concept Check

- Atoms tend to gain or lose electrons until they have obtained an electron configuration that is the same as that of a noble gas.
  - Example: \[ \text{K}^+ (1s^22s^22p^63s^23p^6) \]
    - Lost one electron to obtain electron configuration for \[ \text{Ar} (1s^22s^22p^63s^23p^6) \].
1. Metal atoms containing one, two, or three valence electrons tend to lose electrons to acquire a noble-gas electron configuration. 

<table>
<thead>
<tr>
<th>Group</th>
<th>Charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>IA</td>
<td>1+</td>
</tr>
<tr>
<td>IIA</td>
<td>2+</td>
</tr>
<tr>
<td>IIIA</td>
<td>3+</td>
</tr>
</tbody>
</table>

- charge = group #

2. Nonmetal atoms containing five, six, or seven valence electrons tend to gain electrons to acquire a noble-gas electron configuration.

<table>
<thead>
<tr>
<th>Group</th>
<th>Charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>VIIA</td>
<td>1–</td>
</tr>
<tr>
<td>VIA</td>
<td>2–</td>
</tr>
<tr>
<td>VA</td>
<td>3–</td>
</tr>
</tbody>
</table>

- charge = 8 – group #

3. Elements in Group IVA occupy unique positions relative to the noble gases (could gain or lose four electrons). 

Eg. C and Si

Isoelectronic Species

- A series of ions/atoms containing the same number and configuration of electrons.

- \( \text{O}^2-, \text{F}^-, \text{Ne}, \text{Na}^+, \text{Mg}^{2+}, \text{and Al}^{3+} \)

- \( 1s^22s^22p^6 \)

Concept Check

Choose an alkali metal, an alkaline earth metal, a noble gas, and a halogen so that they constitute an isoelectronic series when the metals and halogen are written as their most stable ions.

- What is the electron configuration for each species?
- Determine the number of electrons for each species.
- Determine the number of protons for each species.
### Section 4.6
#### Lewis Structures for Ionic Compounds

**Formation of an Ionic Compound**
- Ion formation requires the presence of two elements:
  - A **metal** that can donate electrons.
  - A **non-metal** that can accept electrons.
- The electrons lost by the metal are the same ones gained by the nonmetal.
- The positive and negative ions simultaneously formed from such **electron transfer** attract one another.

**Lewis Structure**
- Combination of Lewis symbols that represents either the transfer or the sharing of electrons in chemical bonds.

**The Reaction Between Sodium and Chlorine**

\[
\begin{align*}
\text{Na}^+ + \text{Cl}^- & \rightarrow [\text{Na}]^+ : \text{Cl}^- \rightarrow \text{NaCl} \\
\text{Core} & \left[\text{Ne}\right] \left[\text{Ar}\right]
\end{align*}
\]

**The Reaction Between Sodium and Oxygen**

\[
\begin{align*}
\text{Na}^+ + 2\text{O}^- & \rightarrow [\text{Na}]^+ [\text{O}]^2- \rightarrow \text{Na}_2\text{O}
\end{align*}
\]

**The Reaction Between Calcium and Chlorine**

\[
\begin{align*}
\text{Ca}^{2+} + \text{Cl}^- & \rightarrow [\text{Ca}]^{2+} : \text{Cl}^- \rightarrow \text{CaCl}_2
\end{align*}
\]

**Chemical Formulas for Ionic Compounds**
- Ionic compounds are always neutral; no net charge is present.
- The ratio in which positive and negative ions combine is the ratio that achieves **charge neutrality** for the resulting compound.
- Charges on ions determines the subscripts in the formula
  - Eg. Na\(^{+1}\), O\(^{-2}\) gives Na\(_2\)O
Writing Chemical Formulas for Ionic Compounds

1. The symbol for the positive ions is always written first.
2. The charges on the ions that are present are not shown in the formula.
3. The subscripts in the formula give the combining ratio for the ions.

Example

- Compound formed between \( \text{Li}^+ \) and \( \text{O}^{2-} \)
  - Need two \( \text{Li}^+ \) to balance out the 2- charge on oxygen.
  - Formula is \( \text{Li}_2\text{O} \).

Concept Check

Determine the chemical formula for the compound that is formed when each of the following pairs of ions interact.

- \( \text{Ba}^{2+} \) and \( \text{Cl}^- \)
- \( \text{Fe}^{3+} \) and \( \text{O}^{2-} \)
- \( \text{Pb}^{4+} \) and \( \text{O}^{2-} \)

The Structure of Ionic Compounds

Solid Ionic Compounds (ionic lattices).

- Consists of positive and negative ions arranged in such a way that each ion is surrounded by nearest neighbors of the opposite charge.
- Any given ion is bonded by electrostatic attractions to all the other ions of opposite charge immediately surrounding it.
**Section 4.8**

**The Structure of Ionic Compounds**

**Formula Unit**
- Smallest whole-number repeating ratio of ions present in an ionic compound that results in charge neutrality.
- Chemical formulas for ionic compounds represent the simplest ratio of ions present.
- Eg. Ca$^{2+}$ O$^{2-}$ gives Ca$_2$O$_2^-$ becomes CaO

**Cross-Section of NaCl**

**Section 4.9**

**Recognizing and Naming Binary Ionic Compounds**

**Naming Compounds**
- Binary Compounds:
  - Composed of two elements
  - Ionic and covalent compounds included
- Binary Ionic Compounds:
  - Metal-nonmetal
  - Metal is always present as the positive ion, and the nonmetal is always present as the negative ion.

**Naming Ionic Compounds**

**Examples**
- KCl Potassium chloride
- MgBr$_2$ Magnesium bromide
- CaO Calcium oxide
Section 4.9
Recognizing and Naming Binary Ionic Compounds

Naming Ionic Compounds (for Metals with Variable Charges)

- Metals in these compounds form more than one type of positive charge.
- Charge on the metal ion must be specified.
- Roman numeral indicates the charge of the metal cation (positively charged ion).
- Transition metal cations usually require a Roman numeral.

Examples

<table>
<thead>
<tr>
<th>Compound</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>CuBr</td>
<td>Copper(I) bromide</td>
</tr>
<tr>
<td>FeS</td>
<td>Iron(II) sulfide</td>
</tr>
<tr>
<td>PbO₂</td>
<td>Lead(IV) oxide</td>
</tr>
</tbody>
</table>

Exercise

Name each of the following compounds:

- K₂S
- Fe₂O₃
- CoCl₂

Section 4.9
Recognizing and Naming Binary Ionic Compounds

Metallic Elements with a Fixed Ionic Charge

<table>
<thead>
<tr>
<th>Group I</th>
<th>Group II</th>
<th>Group III</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li⁺, Be²⁺</td>
<td>Na⁺, Mg²⁺</td>
<td>Al³⁺</td>
</tr>
<tr>
<td>K⁺, Ca²⁺</td>
<td>Zn²⁺, Ga³⁺</td>
<td>Ag⁺, Cu²⁺</td>
</tr>
</tbody>
</table>

Exercise

Name each of the following compounds:

- K₂S
- Fe₂O₃
- CoCl₂

Section 4.10
Polyatomic Ions

Polyatomic Ion

- Ion formed from a group of atoms (held together by covalent bonds) through loss or gain of electrons.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>K₂S</td>
<td>potassium sulfide</td>
</tr>
<tr>
<td>Fe₂O₃</td>
<td>iron(III) oxide</td>
</tr>
<tr>
<td>CoCl₂</td>
<td>cobalt(II) chloride</td>
</tr>
</tbody>
</table>
Polyatomic Ions

### Section 4.10

**Polyatomic Ions**

#### Polvatomic Ion

<table>
<thead>
<tr>
<th>Ion</th>
<th>Name</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH₄⁺</td>
<td>ammonium</td>
<td>NH₄⁺</td>
</tr>
<tr>
<td>NO₃⁻</td>
<td>nitrate</td>
<td>NO₃⁻</td>
</tr>
<tr>
<td>NO₂⁻</td>
<td>nitrite</td>
<td>NO₂⁻</td>
</tr>
<tr>
<td>SO₄²⁻</td>
<td>sulfate</td>
<td>SO₄²⁻</td>
</tr>
<tr>
<td>SO₃⁻</td>
<td>sulfite</td>
<td>SO₃⁻</td>
</tr>
<tr>
<td>HSO₄⁻</td>
<td>hydrogen sulfite</td>
<td>HSO₄⁻</td>
</tr>
<tr>
<td>CH⁻</td>
<td>hydroxide</td>
<td>CH⁻</td>
</tr>
<tr>
<td>CN⁻</td>
<td>cyanide</td>
<td>CN⁻</td>
</tr>
<tr>
<td>PO₄³⁻</td>
<td>phosphate</td>
<td>PO₄³⁻</td>
</tr>
<tr>
<td>PO₃⁻</td>
<td>phosphite</td>
<td>PO₃⁻</td>
</tr>
<tr>
<td>HPO₄²⁻</td>
<td>hydrogen phosphate</td>
<td>HPO₄²⁻</td>
</tr>
<tr>
<td>H₂PO₄⁻</td>
<td>phosphite</td>
<td>H₂PO₄⁻</td>
</tr>
</tbody>
</table>

- **Common Polvatomic Ions**

#### Generalizations

1. Most of the polyatomic ions have a negative charge.
2. Two of the negatively charged polyatomic ions, OH⁻ and CN⁻, have names ending in **ide** and the rest of them have names ending in either **ate** or **ite**.

#### Exercise

Which of the following compounds is named incorrectly?

- a) KNO₃ potassium nitrate
- b) TiO₂ titanium(II) oxide
- c) Sn(OH)₄ tin(IV) hydroxide
- d) (NH₄)₂SO₃ ammonium sulfite
- e) CaCrO₄ calcium chromate
Exercise

Which of the following compounds is named incorrectly?

a) KNO$_3$  potassium nitrate
b) TiO$_2$  titanium(II) oxide
c) Sn(OH)$_4$  tin(IV) hydroxide
d) (NH$_4$)$_2$SO$_3$  ammonium sulfite
e) CaCrO$_4$  calcium chromate
titanium(IV) oxide