Chapter 5 Notes: Ions and Ionic Compounds

Sec. 5.1—Simple Ions

1. Relate the electron configuration of an atom to its chemical reactivity.
2. Determine an atom’s number of valence electrons, and use the octet rule to predict what stable ions the atom is likely to form.
3. Explain why the properties of ions differ from those of their parent atoms.

Vocabulary: octet rule, ion, cation, anion

Chemical reactivity—the intensity (amount of energy gained or lost) of chemical reactions partly depends on the electron configuration of the valence shell (outer electron energy level) of the reacting atoms.

The electron configuration for oxygen is: \([O] = 1s^22s^22p^4\)
The electron configuration for neon is: \([\text{Ne}] = 1s^22s^22p^6\)

Magnesium (Mg) will react violently in an atmosphere that contains oxygen but will not react at all when placed in neon (Ne) gas. Notice the 2p orbitals in oxygen have 4 electrons while they have the full number of 6 in neon.

Ne is a noble gas, part of the family in the far right column of the periodic table, labeled Group 18. Each noble gas has a valence shell of electrons that can be labeled \(ns^2np^6\) where \(n\) represents the quantum number of the outer energy level. In most chemical reactions, atoms tend to match the s and p electron configurations of the noble gases, a tendency called the octet rule.

Notice that \(n\) increases by 1 as one goes down one level in the periodic table. The rows of horizontal elements are called periods and \(n\) also represents the row number.

An atom whose outer s and p orbitals do not match the electron configuration of a noble gas will react to lose or gain electrons so the outer orbitals (or valence shell) will be full.

The alkali metals and halogens are the most reactive elements.

- **Alkali metals** are all elements in the far left column in the periodic table, also labeled Group 1. The members of Group 1 only have one electron in their outer energy level. If any element in Group 1 gives up an electron it will achieve the electron configuration of a noble gas and satisfy the octet rule. For potassium (K) the electron configuration change can be written as follows:
  \[ 1s^22s^22p^63s^23p^64s^1 \rightarrow 1s^22s^22p^63s^23p^6 \]

- **The halogens**, located in Group 17, are one column to the left of the noble gases. Group 17 elements need only 1 electron to achieve the electron configuration of a noble gas and satisfy the octet rule. For chlorine (Cl) the electron configuration change can be written:
  \[ 1s^22s^22p^63s^23p^5 \rightarrow 1s^22s^22p^63s^23p^5 \]
Note that in order for potassium and chlorine to achieve a stable configuration, they must lose or gain an electron in order to achieve the electron configuration of argon (Ar).

\[
\begin{align*}
[K^+] &= 1s^22s^22p^63s^23p^6 = [\text{Ne}]3s^23p^6 \quad \text{(potassium ion is a cation)} \\
[\text{Cl}^-] &= 1s^22s^22p^63s^23p^6 = [\text{Ne}]3s^23p^6 \quad \text{(chlorine ion is an anion)} \\
[\text{Ar}] &= 1s^22s^22p^63s^23p^6 = [\text{Ne}]3s^23p^6 \quad \text{(neon is an atom)}
\end{align*}
\]

When an atom loses or gains electrons, it becomes a charged particle called an ion. An ion has a charge, so it forms compounds and also conducts electricity in water. The chemical and physical properties of an ion are completely different from the atom from which it was formed. It is also different from the noble gas that has the same electron configuration. This is because the properties of an atom or its ions depend on the magnetic force created by the electron configuration of the outer shell of electrons.

The Group number corresponds to the number of valence electrons in the particular family of the periodic table. Using this fact and the electron configuration of the noble gases, a shorthand method can be used to write electron configuration. An example of this is shown by the following for the first 18 elements.

<table>
<thead>
<tr>
<th>Atomic No.</th>
<th>Element</th>
<th>Period</th>
<th>Orbitals</th>
<th>Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td>1</td>
<td>s s p</td>
<td>1s(^1)</td>
</tr>
<tr>
<td>2</td>
<td>He</td>
<td>2</td>
<td>1 s p d</td>
<td>1s(^2)</td>
</tr>
<tr>
<td>3</td>
<td>Li</td>
<td>2</td>
<td>2 s 1</td>
<td>[He]2s(^1)</td>
</tr>
<tr>
<td>4</td>
<td>Be</td>
<td>2</td>
<td>2 2</td>
<td>[He]2s(^2)</td>
</tr>
<tr>
<td>5</td>
<td>B</td>
<td>2</td>
<td>2 2 1</td>
<td>[He]2s(^2)2p(^1)</td>
</tr>
<tr>
<td>6</td>
<td>C</td>
<td>2</td>
<td>2 2 2</td>
<td>[He]2s(^2)2p(^2)</td>
</tr>
<tr>
<td>7</td>
<td>N</td>
<td>2</td>
<td>2 2 3</td>
<td>[He]2s(^2)2p(^3)</td>
</tr>
<tr>
<td>8</td>
<td>O</td>
<td>2</td>
<td>2 2 4</td>
<td>[He]2s(^2)2p(^4)</td>
</tr>
<tr>
<td>9</td>
<td>F</td>
<td>2</td>
<td>2 2 5</td>
<td>[He]2s(^2)2p(^5)</td>
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<tr>
<td>10</td>
<td>Ne</td>
<td>2</td>
<td>2 2 6</td>
<td>[He]2s(^2)2p(^6)</td>
</tr>
<tr>
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<td>Na</td>
<td>2</td>
<td>2 2 6 1</td>
<td>[Ne]3s(^1)</td>
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<tr>
<td>12</td>
<td>Mg</td>
<td>2</td>
<td>2 2 6 2</td>
<td>[Ne]3s(^2)</td>
</tr>
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<tr>
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</tr>
<tr>
<td>15</td>
<td>P</td>
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<tr>
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<td>[Ne]3s(^2)3p(^4)</td>
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<tr>
<td>17</td>
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<td>[Ne]3s(^2)3p(^5)</td>
</tr>
<tr>
<td>18</td>
<td>Ar</td>
<td>2</td>
<td>2 2 6 26</td>
<td>[Ne]3s(^2)3p(^6)</td>
</tr>
</tbody>
</table>

Whether an atom loses or gains electrons to become stable and to achieve an outer shell of 8 electrons depends on how much energy is involved. It takes far less energy to remove one electron from the metallic chlorine atom than to add 7 electrons. It takes far less energy to add an electron to the non-metal
fluorine than to take away 7 electrons. A general rule is that metals lose electrons to become cations (positively charged) and non-metals gain electrons to become negatively charged (anions).

**Electron Configuration of Transition Metals:** The transition metals do not follow the octet rule because these elements are filling the d orbitals. Recall that d orbitals require 10 electrons to become filled. Many transition metals can form stable ions with more than one charge.

**Sec. 5.2—Ionic Bonding**

**Vocabulary:** salt, lattice energy, crystal lattice, unit cell

1. Describe the process of forming an ionic bond.
2. Explain how the properties of ionic compounds depend on the nature of ionic bonds.
3. Describe the structure of salt crystals.

**Ionic Bonding**

Because opposite charges attract, anions (negatively charged ions) and cations (positively charged ions) attract each other. An example is table salt, which has the chemical formula NaCl. In table salt, sodium achieves stability by giving up its only valence electron. Chlorine, with seven valence electrons, achieves stability by acquiring one more electron.

The term **salt** is a general term used to describe thousands of different ionic compounds that can form when a cation combines with an anion to form a neutral compound. Anions and cations are held together by magnetic forces.

Because salt crystals are three dimensional, the attraction between anions and cations in a salt is far reaching. In a block of salt, a cation can attract several different anions and an anion can attract several different cations. However the overall composition consists of simple whole-number ratios. In table NaCl, the ratio is 1:1 for one Na⁺ to one Cl⁻.

The formation of salt involves endothermic and exothermic processes. The reaction between solid sodium (Na) and chlorine gas (Cl₂) to form sodium chloride (NaCl) can be broken down into steps as shown in the following figure (see Figure 9 on page 169).

Salts are ionic compounds made of anions and cations. The ratio of cations to anions is always such that an ionic compound has no overall charge. **Ionic compounds** are held together by ionic bonds. In **ionic bonds**, the magnetic force between cations (positive charge) and anions (negative charge). Ionic bonds form between metals (cation) and non-metals (anion). Ionic bonds do not consist of molecules.

The bonds between atoms in **molecular compounds** such as water are different. **Molecules** are held together by covalent bonds. In **covalent bonds**, the atoms are held together by sharing valence electrons. An example is carbon dioxide (CO₂). Covalent bonds are created when non-metals bond to form molecules. Non-metals are molecules as compounds, such as CO₂, and pure elements, such as N₂.
By accounting for all three energies (ionization energy, electron affinity, and lattice energy), we can get a good idea of the energetics involved in such a process.

Metals are pure metallic elements held together by an “electron sea.” A metallic bond forms when mobile valence electrons are shared among metal elements in a stable crystalline structure. Alloys are mixtures of metals. Metals and alloys do not consist of molecules.

The opposite charges of cations and anions attract to form a tightly packed substance of bonded ions called a crystal lattice.

Salts have high melting and boiling points and do not conduct electric current in the solid state, but they do conduct electric current when melted or when dissolved in water.

Salts are made of unit cells that have an ordered packing arrangement.
Sec. 5.3—Names and Formulas of Ionic Compounds

1. Name cations, anions, and ionic compounds.
2. Write chemical formulas for ionic compounds so that an overall neutral charge is maintained.

How do we name ionic compounds?

Most ionic compounds have two word names. The first word is the name of the cation, and the second word is the name of the anion. There is no exception to this rule.

To name ionic compounds, look at the formula and figure out the names of the cation and anion. Then just stick the names together and you've got the name of the compound.

Naming Cations

Use the element name to name cations. For example, the name for the Na⁺ ion is the "sodium." If the cation for an element has more than one charge, put the number charge after the name of the element in Roman numerals. For example, iron can exist as Fe⁺³ or Fe⁺². As a result, you need to specify which cation it is. The Fe⁺² ion has the name “iron (II) and the Fe⁺³ ion has the name “iron (III).”

Naming Anions

Use the first half of the element name and end with an “-ide” to name single anions. Thus, oxygen becomes “oxide,” sulfur becomes “sulfide,” and phosphorus is “phosphide” and so on.

Naming Polyatomic Cations and Anions

Ions with more than one type of element in them are called polyatomic ions. To name a polyatomic ion, look up the correct name in a table. Thus OH⁻ is “hydroxide,” and SO₄²⁻ is “sulfate.”