Modern periodic law

The chemical and physical properties of the elements are periodic functions of their atomic numbers.
Elements are classified as metal, non-metals, or semi-metals, and also fall into groups based on similarities in chemical and physical properties.
The Electronegativity (the power of an atom when in chemical combination to attract electrons to itself) is a periodic property.

<table>
<thead>
<tr>
<th>2.20</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>0.98</td>
<td>1.57</td>
</tr>
<tr>
<td>0.95</td>
<td>1.31</td>
</tr>
<tr>
<td>0.82</td>
<td>1.00</td>
</tr>
<tr>
<td>0.82</td>
<td>0.95</td>
</tr>
<tr>
<td>0.79</td>
<td>0.89</td>
</tr>
</tbody>
</table>

Lewis Structures: Ionic compounds

Whenever possible, the valence electrons in a compound are distributed in such a way that each main-group element in a molecule (except hydrogen) is surrounded by eight electrons (an octet of electrons). Hydrogen should have two electrons in such a structure.

Lewis Structures:

Na · Na⁺ · Ca · Ca²⁺

Sodium atom  Sodium ion  Calcium atom  Calcium ion

**TABLE 3-2**

<table>
<thead>
<tr>
<th>Formulas and Names of Some Monatomic Anions and Cations</th>
</tr>
</thead>
<tbody>
<tr>
<td>F⁻  fluoride ion  H⁺  hydrogen ion</td>
</tr>
<tr>
<td>Cl⁻  chloride ion  Li⁺  lithium ion</td>
</tr>
<tr>
<td>Br⁻  bromide ion  Na⁺  sodium ion</td>
</tr>
<tr>
<td>I⁻  iodide ion  K⁺  potassium ion</td>
</tr>
<tr>
<td>O²⁻  oxide ion  Rb⁺  rubidium ion</td>
</tr>
<tr>
<td>S²⁻  sulfide ion  Cs⁺  cesium ion</td>
</tr>
<tr>
<td>Se³⁻  selenide ion  Be²⁺  beryllium ion</td>
</tr>
<tr>
<td>Te⁴⁻  telluride ion  Ca²⁺  calcium ion</td>
</tr>
<tr>
<td>N³⁻  nitride ion  Sr²⁺  strontium ion</td>
</tr>
<tr>
<td>P⁵⁻  phosphide ion  Ba²⁺  barium ion</td>
</tr>
<tr>
<td>As³⁻  arsenide ion  B³⁺  boron ion</td>
</tr>
<tr>
<td>Sb³⁻  antimonide ion  Al³⁺  aluminum ion</td>
</tr>
<tr>
<td>H⁻  hydride ion  Ga³⁺  gallium ion</td>
</tr>
</tbody>
</table>
Lewis Structures: Covalent Compounds

Whenever possible, the valence electrons in a compound are distributed in such a way that each main-group element in a molecule (except hydrogen) is surrounded by eight electrons (an octet of electrons). Hydrogen should have two electrons in such a structure.

Ammonia (NH₃)

Nitrogen (Group V)
5 valence
+ 2 in 1s = 7 electrons

Lewis Structures
Drawing Lewis Structures

1. Count up the total number of valence electrons available (symbolized by $A$) by first adding the group numbers of all the atoms present. If the species is a negative ion, add the absolute value of the total charge; if it is a positive ion, subtract it.

2. Calculate the total number of electrons needed ($N$) for each atom to have its own noble-gas set of electrons around it (two for hydrogen, eight for the elements from carbon on in the periodic table).

3. Subtract the number in step 1 from the number in step 2. This is the number of shared (or bonding) electrons present ($S$).

4. Assign two bonding electrons (as one shared pair) to each connection between two atoms in the molecule or ion.

5. If any of the electrons earmarked for sharing remain, assign them in pairs by making some of the bonds double or triple bonds. In some cases, there may be more than one way to do this. In general, double bonds form only between atoms of carbon, nitrogen, oxygen, and sulfur.

6. Assign the remaining electrons as lone pairs to the atoms, giving octets to all atoms except hydrogen.

7. Determine the formal charge on each atom, and write it next to that atom. Check that the formal charges add to give a correct total charge on the molecule or molecular ion.

Formal charge = group number - lone-pair electrons - $\frac{1}{2}$ (number of electrons in bonding pairs)

C
\[ \begin{align*}
&\text{H} \\
&\text{H--C--H} \\
&\text{H}
\end{align*} \]
Drawing Lewis Structures

3. Subtract the number in step 1 from the number in step 2. This is the number of shared (or bonding) electrons present (S).

4. Assign two bonding electrons (as one shared pair) to each connection between two atoms in the molecule or ion.

Drawing Lewis Structures

1. Count up the total number of valence electrons available (symbolized by A) by first adding the group numbers of all the atoms present. If the species is a negative ion, add the absolute value of the total charge; if it is a positive ion, subtract it.

2. Calculate the total number of electrons needed (N) for each atom to have its own noble-gas set of electrons around it (two for hydrogen, eight for the elements from carbon on in the periodic table).
Drawing Lewis Structures

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Drawing Lewis Structures

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\[
\text{C} \equiv \text{O}
\]

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\[
\text{C} \equiv \text{O}
\]

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\[
\text{C} \equiv \text{O}
\]
6. Assign the remaining electrons as lone pairs to the atoms, giving octets to all atoms except hydrogen.

C≡O

7. Determine the formal charge on each atom, and write it next to that atom. Check that the formal charges add to give a correct total charge on the molecule or molecular ion.

H
H-C-H
H

H
H-C-H
H
Drawing Lewis Structures

7. Determine the \textit{formal charge} on each atom, and write it next to that atom. Check that the formal charges add to give a correct total charge on the molecule or molecular ion.

\[
\begin{align*}
\text{H} & \quad \text{C} & \quad \text{H} \\
\text{H} & \quad \text{C} & \quad \text{H}
\end{align*}
\]

Naming Compounds in Which Covalent Bonding Occurs

<table>
<thead>
<tr>
<th>TABLE 3-4</th>
<th>Prefixes Used for Naming Binary Molecular Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>Number</td>
<td>Prefix</td>
</tr>
<tr>
<td>1</td>
<td>mono-</td>
</tr>
<tr>
<td>2</td>
<td>di-</td>
</tr>
<tr>
<td>3</td>
<td>tri-</td>
</tr>
<tr>
<td>4</td>
<td>tetra-</td>
</tr>
<tr>
<td>5</td>
<td>penta-</td>
</tr>
<tr>
<td>6</td>
<td>hexa-</td>
</tr>
<tr>
<td>7</td>
<td>hepta-</td>
</tr>
<tr>
<td>8</td>
<td>octa-</td>
</tr>
<tr>
<td>9</td>
<td>nona-</td>
</tr>
<tr>
<td>10</td>
<td>deca-</td>
</tr>
<tr>
<td>11</td>
<td>undeca-</td>
</tr>
<tr>
<td>12</td>
<td>dodeca-</td>
</tr>
</tbody>
</table>

Blue = Nitrogen
Red = Oxygen

\[
\begin{align*}
\text{a.)} & \\
\text{b.)} & \\
\text{c.)} & \\
\text{d.)} & \\
\text{e.)} & \\
\text{f.)} &
\end{align*}
\]
Naming Compounds that Contain Polyatomic Ions

3-7 the Shapes of Molecules

The VESPR theory

**Steric Number**

\[ SN = (\text{number of atoms bonded to a central atom}) + (\text{number of lone pairs on central atom}) \]
The VSEPR Theory

The Valence Shell Electron-Pair Repulsion Theory

Electron pairs in the valence shell of an atom repel each other on a spherical surface formed by the underlying core of the atom.

The geometry which applies to a particular arrangement is determined by the **steric number** (SN) of the central atom.

“Steric” means “having to do with space.” The steric number of an atom in a molecule can be determined by drawing the Lewis structure of the molecule and adding the number of atoms that are bonded to it and the number of lone pairs that it has.

Geometry and Steric Number

- SN = 2 Linear 180°
- SN = 3 Trigonal planar 120°
- SN = 4 Tetrahedral 109.5°
- SN = 5 Trigonal bipyramidal
  - 90° (equatorial - axial or
  - 120° (equatorial – equatorial)
- SN = 6 Octahedral 90°
- SN for Double and triple bonds count the same as single bonded atoms
- SN = 5 Lone pairs occupy equatorial positions in preference to axial positions.
- When lone pairs are present, the situation is more complicated due to repulsive forces
  - Lone pr. vs lone pair >
  - Lone pr vs. bonding pair >
  - Bonding pr. Vs. bonding pair

Examples with no Lone Pairs on the central Atom

The VSEPR Theory
The VSEPR Theory

(a) PF₅
(b) SF₄
(c) SF₆
(d) ClF₃
(e) XeF₂

Axial
Equatorial

107.3°
104.5°

Trigonal bipyramid
Distorted tetrahedron
Linear

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Dipole Moments

- Bonded atoms share electrons unequally, whenever they differ in Electronegativity
- E.g., HCl. The Cl atom carries a slightly negative electric charge and the H atom a slightly positive charge of equal magnitude. Aligns itself in an electric field
- Dipolar or polar molecules posses a dipole moment, $\mu$
- To find $\mu$, first draw the VESPR molecular geometry.
  - Next assign a dipole moment to each bond,
  - then vectorially add the dipole moments for each bond
3-8 Elements Forming more than One Ion

- Oxidation State (Oxidation Number)
  - Ionic vs Covalent bonding
  - **Not** formal electric charges, rather what the charge would be if the compound were ionic
  - Range from -3 to +7

- Examples
  1. \( \text{CrO}_3 \)
  2. \( \text{TiCl}_3 \)
  3. \( \text{Mn}_3\text{N}_2 \)
  4. \( \text{VCl}_4 \)
  5. \( \text{Mn}_2\text{O}_7 \)

---

**Chapter 3**

Chemical Periodicity and the Formation of Simple Compounds

- Examples / exercises
  - All (3-1 to 3-16)

- HW Problems
  - 8, 9, 18, 34, 45, 57, 59, 60, 69, 70