Section 9.1: interesting bit about HIV and drugs that inhibit HIV protease (pro-tee-ace)

Section 9.2: How are electrons distributed in chemical bonds?

<table>
<thead>
<tr>
<th>Types of Atoms</th>
<th>Type of Bond</th>
<th>Characteristic of Bond</th>
</tr>
</thead>
<tbody>
<tr>
<td>Metal and nonmetal</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Nonmetal and nonmetal</td>
<td></td>
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<tr>
<td>Metal and metal</td>
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</tbody>
</table>

Lewis symbols for atoms
- Determine the valence electrons for a N atom: ______________________________________________
- Determine the total number of valence electrons _________________________
- Lewis symbols are valence e\(^{-}\) shown as dots around the element’s symbol
- Draw as single dots (up to 4) then pair up
- Help to identify bonding in simple molecules

Draw Lewis symbols for:

Li Na K B C N O F Ne

Section 9.3: How can chemists represent structures of molecules (Lewis structures)?

Electron sea model

Section 9.4: Lewis symbols for ionic compounds (transferred electrons)
- Electrons are transferred in ionic compounds
- Draw cation and anion separately
- Ions are drawn with square brackets and charges outside of the brackets
- Ions gain or lose electrons so that they have 8 valence electrons (main groups only)
- An atom or ion with 8 electrons is said to have an “Octet”
- Draw the Lewis structures for MgO and MgCl\(_2\).
The Octet

Octet Guideline: an atom tends to gain, lose, or share

- **Caution:**
- **Atoms:**

Why: ____________________________________________________________

**Section 9.5: Covalent Bonding**
- If 2 similar, non-metallic atoms bond, ____________________________________________________________.
- atoms share pairs of e\(^{-}\) ________________________________________________________________
- 1 pair shared e\(^{-}\) equals _________________________________________________________________

Example: ______________________________________________________________

Video of bonding in H\(_2\):

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**Lewis structures: dots or lines?**
In Lewis structures, use *dots* or lines to show electron pairs in covalent bonds

**Section 9.7: How to draw Lewis structures for molecules:**
1. Add up all valence e\(^{-}\) from the atoms involved, including charge if an ion
2. Arrange atoms as bonded
   - Look at formula, (many around 1)
   - What’s bonded to what (H–O in oxo acids)
   - “typical” # of bonds
3. Place 2 e\(^{-}\) between each bonded atom
4. Add in e\(^{-}\) pairs to make octets for outer atoms (not H)
5. Place extra e\(^{-}\) in pairs around central atom
6. If no octet on central atom, share e\(^{-}\) to make multiple bonds
7. If all else fails, LOOK UP STRUCTURE
8. Large atoms more likely to violate octet rule (3d orbitals for S, P, Cl)

Additional tips:
- The *typical* number of bonds for an atom: 8 − group number = bonds
- Carbon atoms usually bonded together in a chain
- Don’t make too many double bonds (never for H, B, or F)
Examples: NH$_3$  BH$_3$  C$_2$Cl$_4$  HCO$_2$H  NaCN

Practice:
CH$_4$  PF$_3$  CH$_3$CH$_2$OH  CH$_3$CO$_2$H  XeF$_4$  C$_2$F$_2$  AsF$_6^-$

Section 9.9: Exceptions to the octet guideline:
- Small atoms (H through B)
- Atoms that can access the next d orbital set
  3$^{\text{rd}}$ row: S, P, Cl
  4$^{\text{th}}$ row: As, Se, Br
  5$^{\text{th}}$ row: Sb, Te, I, Xe
- Called “expanded octets”

Multiple Bonds
- 1 shared pair of e$^-$
- 2 pairs of e$^-$
- 3 pairs of e$^-$
- Covalent bonds have shared e$^-$, so for e$^-$ to be shared,
- bond distances:
- Rank the following from longest bond to shortest: O$_2$  H$_2$  N$_2$
Section 9.6 Bond Polarity and Electronegativity

- In covalent bonds, $e^-$ are shared, but not always equally.
- Unequal sharing of $e^-$ results in __________________________ bonds.
- **Electronegativity:**

How does EN change:
- L to R across a period (row)? _________________________________________________
- down a group (column)? _____________________________________________________

Be able to identify these 3 bond types from differences in EN values

The Continuum of Bond Types

Dipole Moments

- **Dipole:** two different “ends” of the molecule
- **Dipole moment** is term for size of dipole

<table>
<thead>
<tr>
<th>Table 9.2 Dipole Moments of Several Molecules in the Gas Phase</th>
</tr>
</thead>
<tbody>
<tr>
<td>Molecule</td>
</tr>
<tr>
<td>Cl$_2$</td>
</tr>
<tr>
<td>CIF</td>
</tr>
<tr>
<td>HF</td>
</tr>
<tr>
<td>LiF</td>
</tr>
</tbody>
</table>

Determine the bonding type for the following: (ionic, polar covalent, or covalent)

1. N–F
2. C–F
3. C–H
4. Li–Cl

Which of the bonds is the most polar?
Section 9.8: Resonance and Formal Charge

For some molecules, more than one Lewis structure is possible. Draw the Lewis structure for thiocyanate, $\text{NCS}^-$

Resonance structure isn’t sometimes one, sometimes the other. It’s always a _________________________

Formal charge is a way to determine best structure

\[ FC = (\text{valence } e^-) - (\text{# of bonds}) - (\text{non-bonded } e^-) \] (for each atom)

Best Lewis structure is the one with:
- FC on each atom is _______________________________________________
- negative FCs are on the ____________________________________________

Formal charges are not real

In resonance structures, electrons are ____________________________________ (not confined to one area)