4.1 Lewis Theory of Bonding

The key ideas of the Lewis theory of bonding are:

- Atoms and ions are stable if they have a noble gas-like electron structure; i.e., a stable octet of electrons.
- Electrons are most stable when they are paired.
- Atoms form chemical bonds to achieve a stable octet of electrons.
- A stable octet may be achieved by an exchange of electrons between metal and nonmetal atoms.
- A stable octet of electrons may be achieved by the sharing of electrons between nonmetal atoms.
- The sharing of electrons results in a covalent bond.

Rules for Drawing Lewis Structures

Step 1 Use the last digit of the group number from the periodic table to determine the number of valence electrons for each atom.

Step 2 Place one electron on each of the four sides of an imaginary rectangle enclosing the central atom before pairing any electrons.

Step 3 If there are more than four valence electrons, pair up the electrons as required placing all of the valence electrons.

Step 4 Use the unpaired electrons to bond additional atoms with unpaired electrons to the central atom until a stable octet is obtained.

Procedure for Drawing Lewis Structures

Step 1 Arrange atoms symmetrically around the central atom (usually listed first in the formula, not usually oxygen and never hydrogen).

Step 2 Count the number of valence electrons of all atoms. For polyatomic ions, add electrons corresponding to the negative charge, and subtract electrons corresponding to the positive charge on the ion.

Step 3 Place a bonding pair of electrons between the central atom and each of the surrounding atoms.

Step 4 Complete the octets of the surrounding atoms using lone pairs of electrons. (Remember hydrogen is an exception.) Any remaining electrons go on the central atom.

Step 5 If the central atom does not have an octet, move lone pairs from the surrounding atoms to form double or triple bonds until the central atom has a complete octet.

Step 6 Draw the Lewis structure and enclose polyatomic ions within square brackets showing the ion charge.

4.1 (additional notes)

Co-ordinate Covalent Bond:

- A covalent bond in which one atom donates both bonding electrons. (This is an exception to regular covalent bonding where one electron is donated by each atom.)

Examples:

Name	Formula	Electron dot diagram
Ammonium	NH_4^{+1}	
Hydronium	H_3O^{+1}	
Nitrosyl trifluoride	NF_3O	

Resonance Structures:

- When a structure containing a double bond can be drawn with the double bond in two or more locations without changing the arrangement of atoms, a resonance structure is said to exist. These different structures for the same molecule are called “resonance hybrids”.
- Example: SO_3 (sulfur trioxide)

- The bond lengths and strengths have been experimentally determined to fall between single and double bonds.
- The electrons forming the double bond(s) are said to be “delocalized” (shared) over all the bonds.

Exceptions to the Octet Rule:

1. Example: PCl_5

- Molecules that have more than four atoms bonded to the central atom.

2. Example: BF_3

- Molecules that contain no double bonds and whose central atom has fewer than four bonding electrons.

3. Example: NO

- Molecules containing an odd number of electrons.
- Molecules of this type are also “paramagnetic” because they are attracted by a magnetic field.

Homework
- Practice 1,3,4,5,10,11,12
- Questions 2,3,4,5