Unit 1, Lesson 05: Introduction to Bonding
(pages 162-166)

- atoms are most stable (lowest energy) when they have a full outer electron shell (a stable octet electron arrangement, s^2p^6)
- the chemical behaviour of atoms results from atoms losing, gaining or sharing valence electrons to obtain a more stable electron arrangement
- the ease with which the valence electrons are lost (ionization energy, IE) or gained (electron affinity, electronegativity, EN) determines what kinds of bonds and compounds the elements will form
- the Noble gases have a stable octet; it takes too much energy to remove a valence electron (high ionization energy, low electronegativity and low electron affinity), so they usually do not form bonds

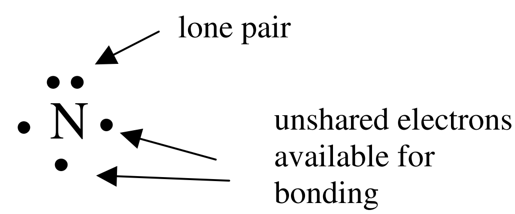
There are three types of bonding, based on the ionization energies, electronegativities and electron affinities of the atoms involved:
- ionic bonding, usually between metals (low IE, low EN) and non-metals (high IE, high EN)
- metallic bonding, usually between metals (low IE, low EN) and metals (low IE, low EN)
- covalent bonding, usually between non-metals (high IE, high EN) and non-metals (high IE, high EN)

We will use electron dot (Lewis dot) diagrams to represent bonding.

Review of Electron Dot (Lewis) Diagrams
(page 163)

Because it is the valence electrons that are usually involved in bonding, we can represent bonding using electron dot (Lewis) diagrams. Remember, when drawing electron dot diagrams:
- write the symbol for the element and show only the outer (valence) electrons
- the number of valence electrons is equal to the Group Number (in Roman Numerals) on the Periodic Table for each element
- follow the convention of only doubling up the electrons after all four "orbitals" have one electron each

eg. the electron dot diagram for nitrogen (Group V) would be:



eg. the electron dot diagrams for the elements of the third period would be:

Group I	Group II	Group III	Group IV	Group V	Group VI	Group VII	Group VIII
Na	Mg	Al	Si	P	S	Cl	Ar

It is important to draw correct Lewis diagrams in order to determine how many unshared electrons an atom has because:
- orbitals that contain two electrons are full, so they do not usually participate in bonding. Electrons in full orbitals are called "lone pairs"
- orbitals that contain only one electron will form bonds in order to obtain an additional electron to fill the orbital, so it is the unshared electrons that usually participate in bonding