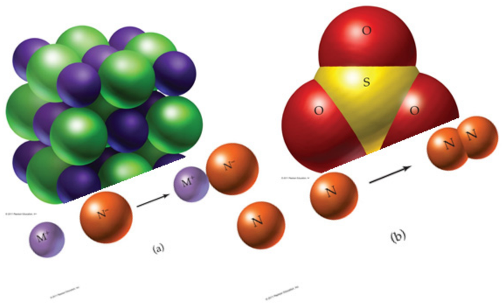
Chapter 12 – Chemical Bonding



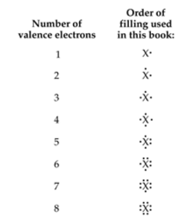
1916 – Gilbert Newton Lewis develops theory of chemical bonding

Notes that noble gases are unusually stable.

All have eight ***valence electrons*** (except helium which has two)

This is known as the **octet rule**.

Atoms achieve noble gas electron configuration by **transferring** or **sharing electrons** between them and in the process producing compounds.



*Valence electrons are outermost energy level electrons. Core electrons are the inner electrons.

Ionic Bonding

Atoms achieve noble gas electron configuration by transferring one or more electrons.

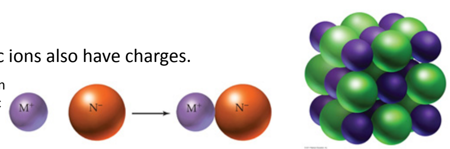
Metals lose electrons and become positively charged **cations**.

Nonmetals gain electrons and become negatively charged **anions**.

The basic particle of an ionic compound is known as a **formula unit**.

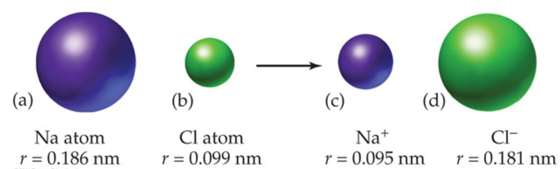
Polyatomic ions also have charges.

M^+ = metallic ion
N^- = nonmetallic ion



Anions are always larger than the parent atom from which they came.

Cations are always smaller than the parent atom from which they came.

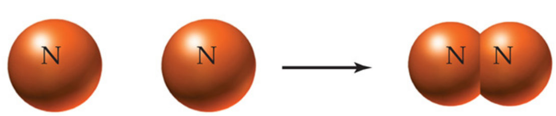


Covalent Bonding

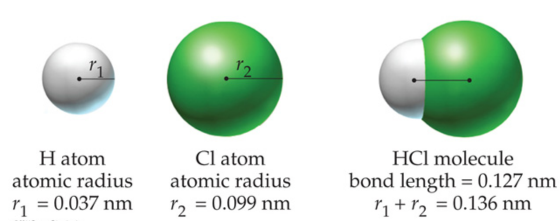
Atoms achieve noble gas electron configuration by sharing one or more electrons.

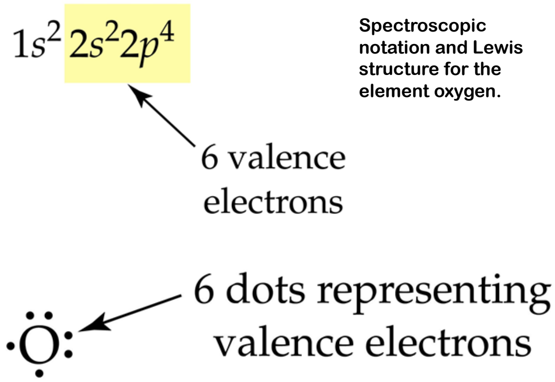
This occurs primarily between nonmetal atoms.

The basic particle of an covalent compound is known as a **molecule**.



For covalent compounds, the internuclear distance is not quite the sum of the radii of the elements that are bonding because there is some overlap of the electron clouds of the atoms.

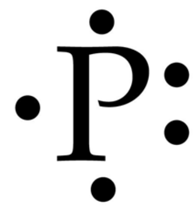




Spectroscopic notation and Lewis structure for the element oxygen.

Question:

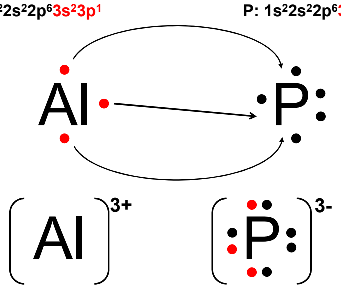
Draw the Lewis structure for phosphorus. What are its valence electrons?



The valence electrons are $3s^2 3p^3$. Note that the Lewis structure does not differentiate between s and p sublevel electrons.

Example: Phosphorus as part of an ionic compound (AlP) vs. phosphorus as part of a covalent compound (PCl_3).

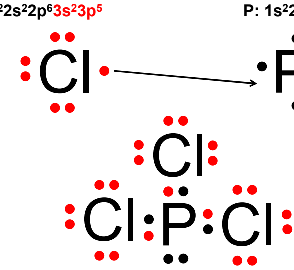
Al: $1s^2 2s^2 2p^6 3s^2 3p^1$ P: $1s^2 2s^2 2p^6 3s^2 3p^3$



Result: Aluminum = Octet like previous noble gas Ne; Phosphorus = Octet like next noble gas Ar.

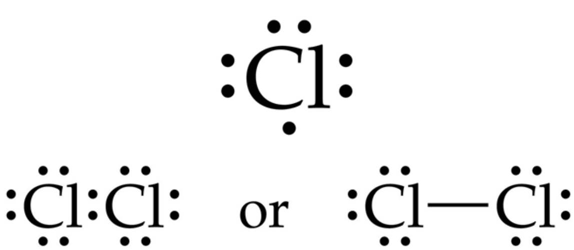
Example: Phosphorus as part of an ionic compound (AlP) vs. phosphorus as part of a covalent compound (PCl_3).

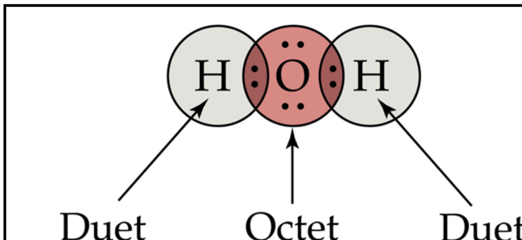
Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$ P: $1s^2 2s^2 2p^6 3s^2 3p^3$

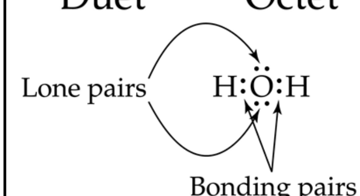


Result: Chlorine(s) = Octet like next noble gas Ar; Phosphorus = Octet like next noble gas Ar.

Example: Show the formation of diatomic chlorine using Lewis Structures.



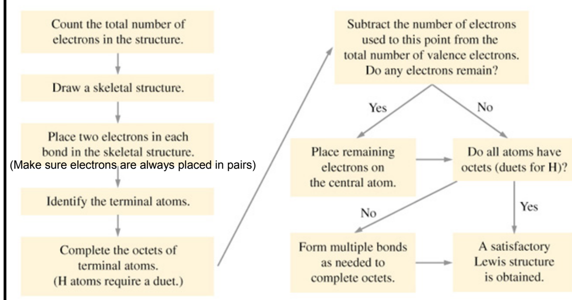




Important terminology you should know

Important Note: Lewis structures cannot be used to determine the spatial geometry of a molecule. H_2O is NOT linear!

General Procedure for Writing Lewis Structures for Molecules



Example:

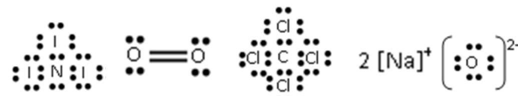
Draw the most likely Lewis structure for the following:

a. Nitrogen triiodide

b. Diatomic oxygen

c. Carbon tetrachloride

d. Sodium oxide



Note: The textbook refers to the use of lines as the "structural formula" vs. the "electron dot formula" when only dots are used.

Polyatomic Ions and Resonance

When determining the electron dot structure for a polyatomic ion, the number of deficient or surplus electrons must be taken into account.

The overall charge on the polyatomic ion determines the number of electrons to be added or subtracted.

In a **resonance structure**, more than one legitimate, **equivalent** structure is possible. This is common to, but not exclusive of polyatomic ions.

For molecules or ions that exhibit resonance, the true structure is a "hybrid" of the different structures.

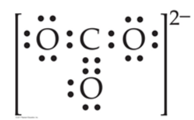
Example: Determine a reasonable Lewis Dot Structure for the carbonate polyatomic ion.

Answer: The carbonate polyatomic ion is CO_3^{2-}.

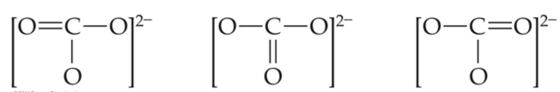
Valence electron available:

C	=	4
O	=	6 x 3
Charge	=	2 extra electrons (2-)

Total electrons = 24



The double bond allows an octet without violating the number of electrons available.



The double bond can be placed equivalently on any one of the three carbon-oxygen bonds. The carbonate polyatomic ion exhibits resonance. (Note: Nonbonding electron pairs have been omitted from the image for clarity.)

On Your Own:

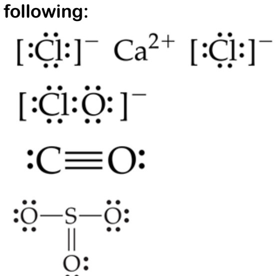
Write the most likely Lewis Structures and determine if resonance is present for the following:

Calcium Chloride

Hypochlorite ion (ClO^-)

Carbon monoxide

Sulfur trioxide

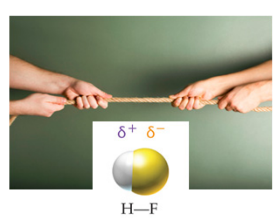


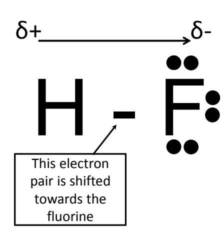
Sulfur trioxide exhibits resonance.

Bond Polarity

Caused by the "unequal sharing" of electrons in a bond between two atoms.

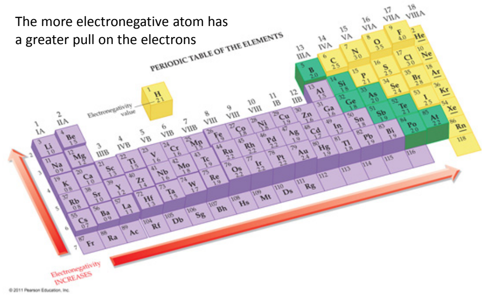
Like a "tug-o-war" between the two bonding atoms.





Bond Polarity

The strength of the "dipole" (i.e. how polar the bond is) is determined by the difference in electronegativity between two atoms, ΔEN



Bond Character

The type of bond (bond character) is determined by ΔEN

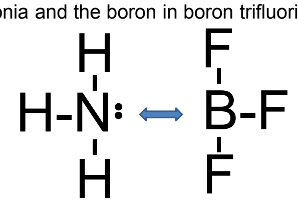
ΔEN	**Bond Character (Bond type)**
0	Pure (nonpolar) covalent
> 0 to 1.7	Polar covalent
> 1.7	Ionic

The transition between polar covalent and ionic is not sudden and values vary slightly from one resource to another.

Coordinate Covalent Bonds

A coordinate covalent bond is one in which the electrons shared between two atoms all come from only one of the atoms.

An example is the bond that forms between the nitrogen in ammonia and the boron in boron trifluoride.

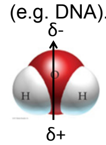


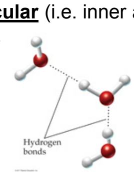
Hydrogen Bonding

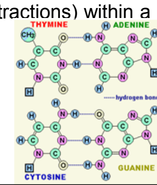
A particularly strong polarity occurs when a molecule has **hydrogen bonded to nitrogen, oxygen or fluorine**: H-N, H-O or H-F

Since a polar molecule acts like a tiny electrical "magnet", it produces strong **intermolecular forces** between atoms.

This can not only affect things like melting and boiling point (e.g. H_2O) but also the "shape" of a molecule through **intramolecular** (i.e. inner attractions) within a molecule (e.g. DNA).





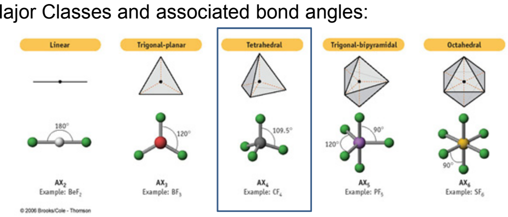


VSEPR Theory

Valence shell electron pair repulsion theory: Electron pairs around the central atom geometrically arrange themselves to minimize repulsions (i.e. as far away from each other as possible). New, **hybridized orbitals** are formed.

Determine the geometry of the molecule with respect to a central atom.

Major Classes and associated bond angles:

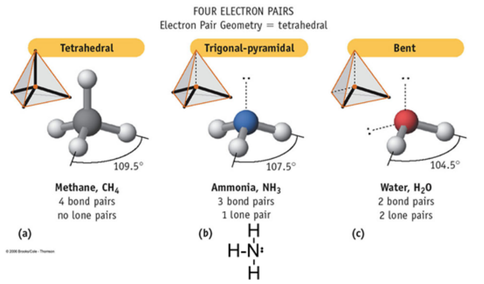


- Geometrically, treat multiple bonds in a dot structure as though they were single bonds.

- Be able to differentiate between **electron geometry** (always a major class) and **molecular geometry** (major class or subclass)

- Lone pairs on the central atom will affect the geometry of the molecule.

Major Class and Sub-Classes of Tetrahedral



Notice that water (example c) has an electronic geometry that is tetrahedral but a molecular geometry that is bent.

Molecular Geometry Classes and Subclasses

TABLE 9.2 Electron-Pair Geometries and Molecular Shapes for Molecules with Two, Three, and Four Electron Domains About the Central Atom

Total Electron Domains	Electron-Domain Geometry	Bonding Domains	Non-bonding Domains	Molecular Geometry	Example
2 pairs	Linear	2	0	Linear	O=C=O
3 pairs	Trigonal planar	3	0	Trigonal planar	BF_3
		2	1	Bent	$[NO_2]^-$
4 pairs	Tetrahedral	4	0	Tetrahedral	CH_4
		3	1	Trigonal pyramidal	NH_3
		2	2	Bent	H_2O

TABLE 9.3 Electron Pair Geometries and Molecular Shapes for Molecules with Five and Six Electron Pairs Domains About the Central Atom

Number of Electron Domains	Electron-Domain Geometry	Bonding Domains	Nonbonding Domains	Molecular Geometry	Example
5 domains	Trigonal bipyramidal	5	0	Trigonal bipyramidal	PCl_5
		4	1	Seesaw	SF_4
		3	2	T-shaped	ClF_3
		2	3	Linear	XeF_2
6 domains	Octahedral	6	0	Octahedral	SF_6
		5	1	Square pyramidal	BrF_5
		4	2	Square planar	XeF_4

Polarity Revisited

Only two things can make a molecule nonpolar:

1. A ΔEN value of 0 between atoms

2. A molecular symmetry that causes the dipoles (polarity) to cancel out.