Chemical Bonding

Bonding

Bond – force that holds groups of atoms together and makes them function as a unit.

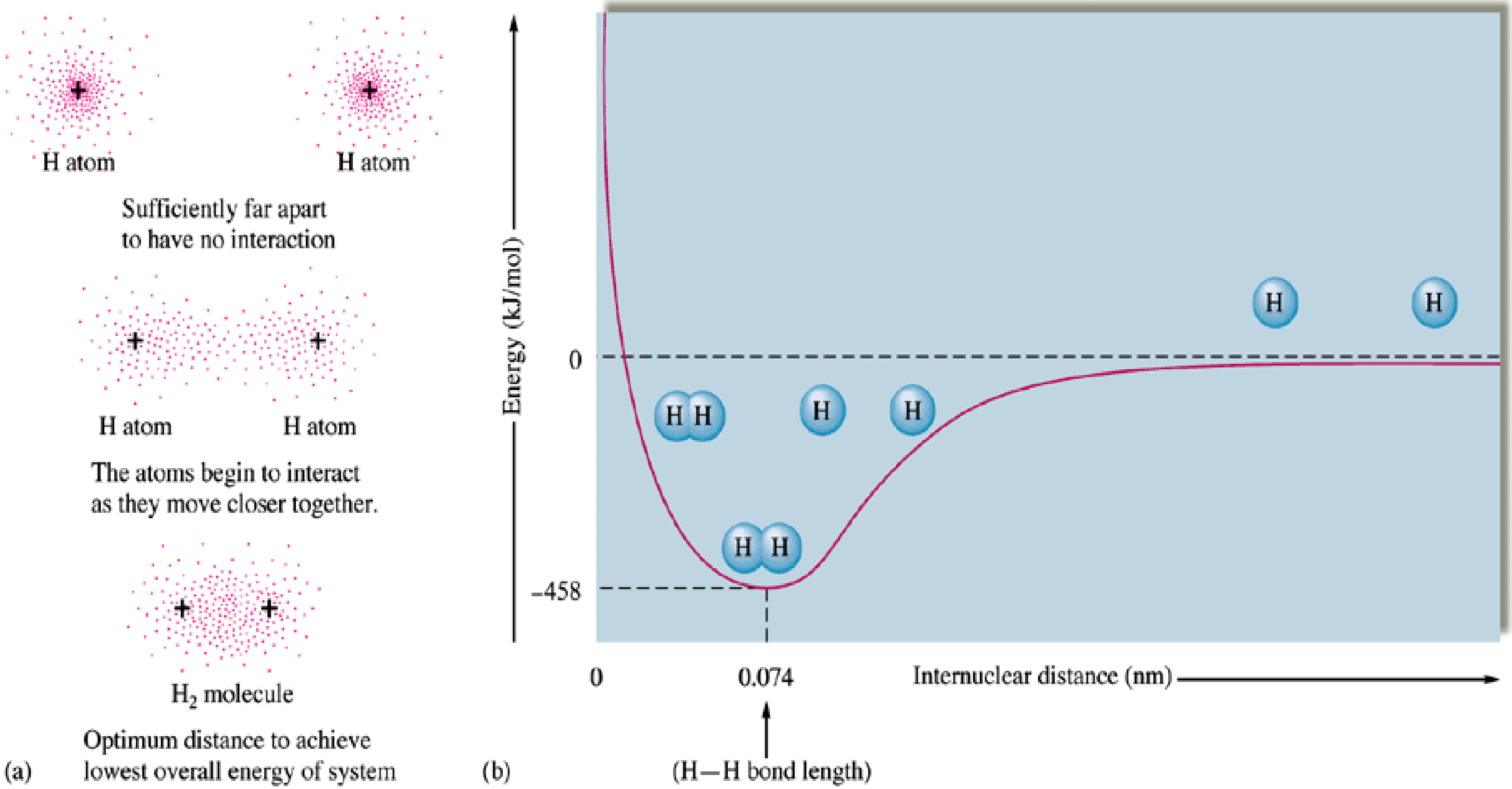
- There is a limit to how close the atoms can get.
- The most stable distance is affected by electron repulsion and attraction between the two nuclei

Bond energy – energy required to break a chemical bond. Bonds will be created if it allows the system (two or more atoms) to achieve the lowest possible energy state.

Usually when a chemical bond is

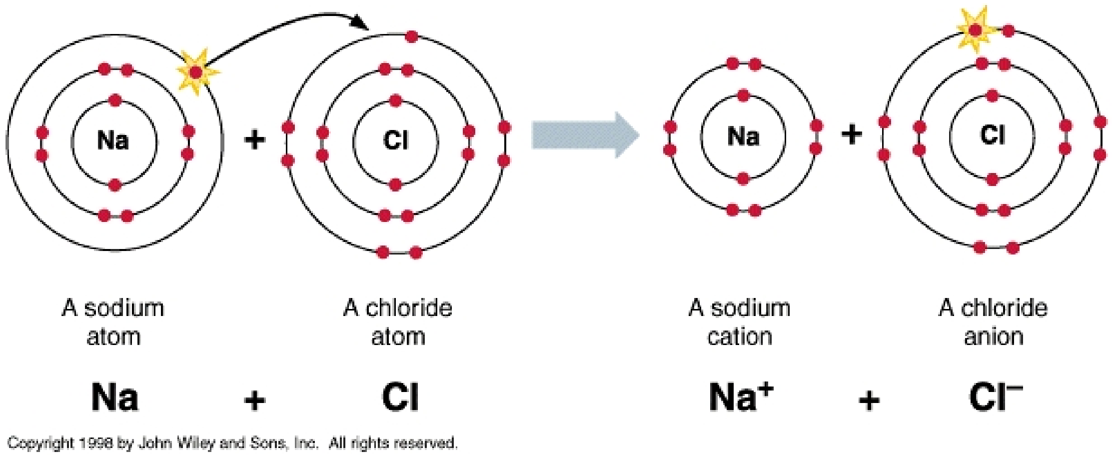
- **Formed**
  energy is **released** (overall energy change): **EXOTHERMIC REACTION**
- **Broken**
  energy is **absorbed** (overall energy change): **ENDOTHERMIC REACTION**

Bond Energy

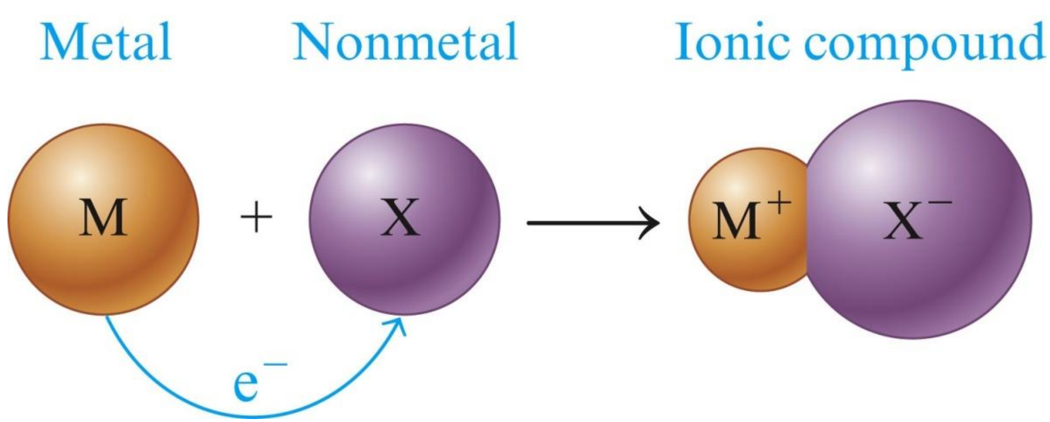

H atom
H atom
Sufficiently far apart
to have no interaction
H atom
H atom
The atoms begin to interact
as they move closer together.
H_2 molecule
Optimum distance to achieve
lowest overall energy of system
(a)
Energy (kJ/mol)
0
–458
H H
H
H
H
H
H H
0
0.074
Internuclear distance (nm)
(H—H bond length)
(b)

Bonding

Ionic Bonding: **results from the attraction between oppositely charged ions**

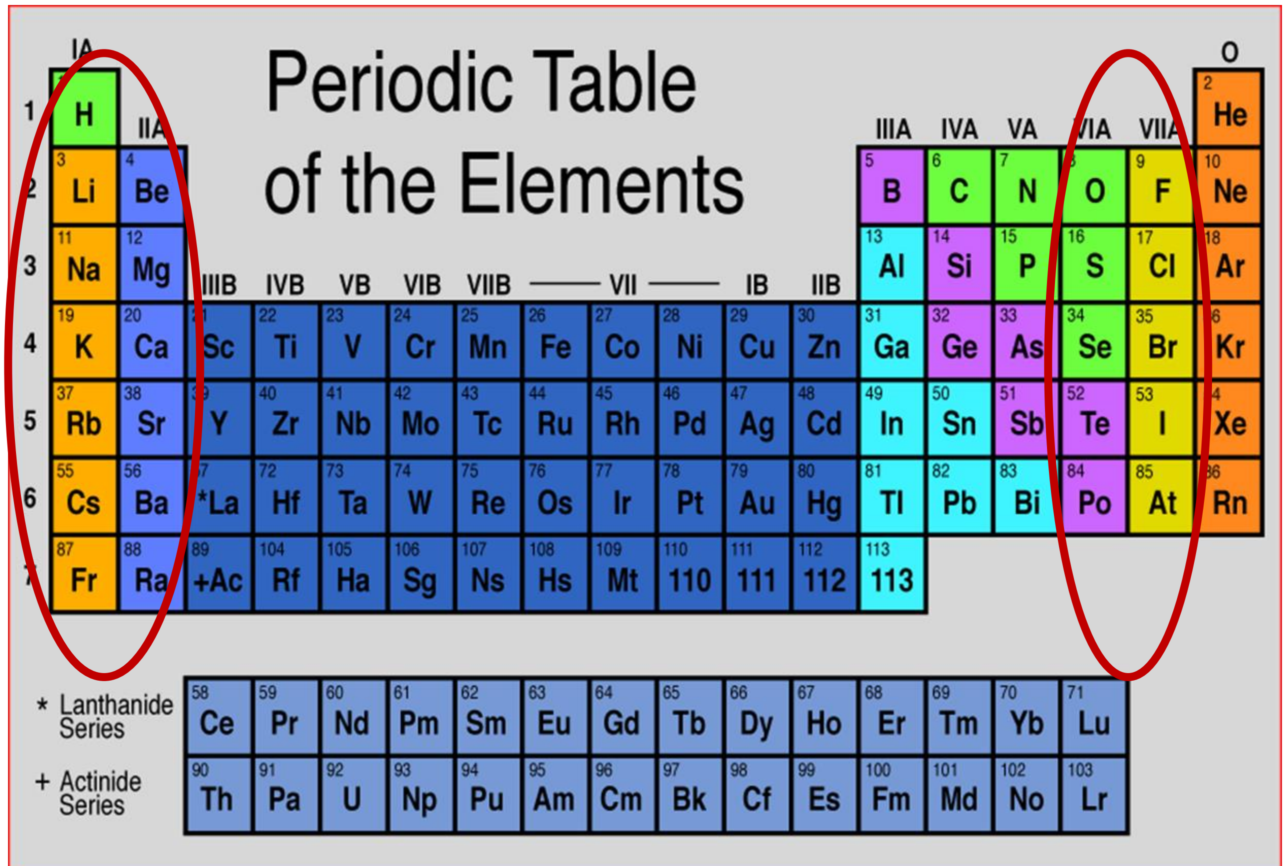


- **Ionic compound results when a metal reacts with a nonmetal.**



Bonding

Best examples of ionic bonding between Groups IA, IIA and VIA, VIIA



Characteristics of Ions and Ionic Compounds

Stable Electron Configurations and Charges on Ions

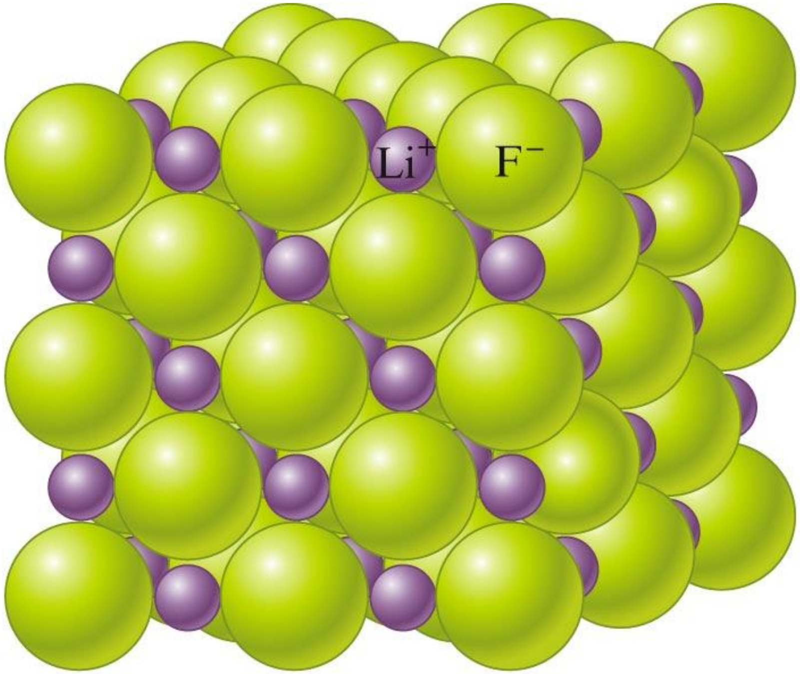
- **Valence electrons involved in chemical bonding.**
- **For most atoms this means a total of eight valence electrons**
- **Octet Rule: When forming chemical bonds, atoms take on an electron configuration of the nearest noble gas**
- **Ionic Bonds- electrons are transferred from the outer shell of one atom to the outer shell of another atom to form the octet**

Characteristics of Ions and Ionic Compounds

B. Ionic Bonding and Structures of Ionic Compounds

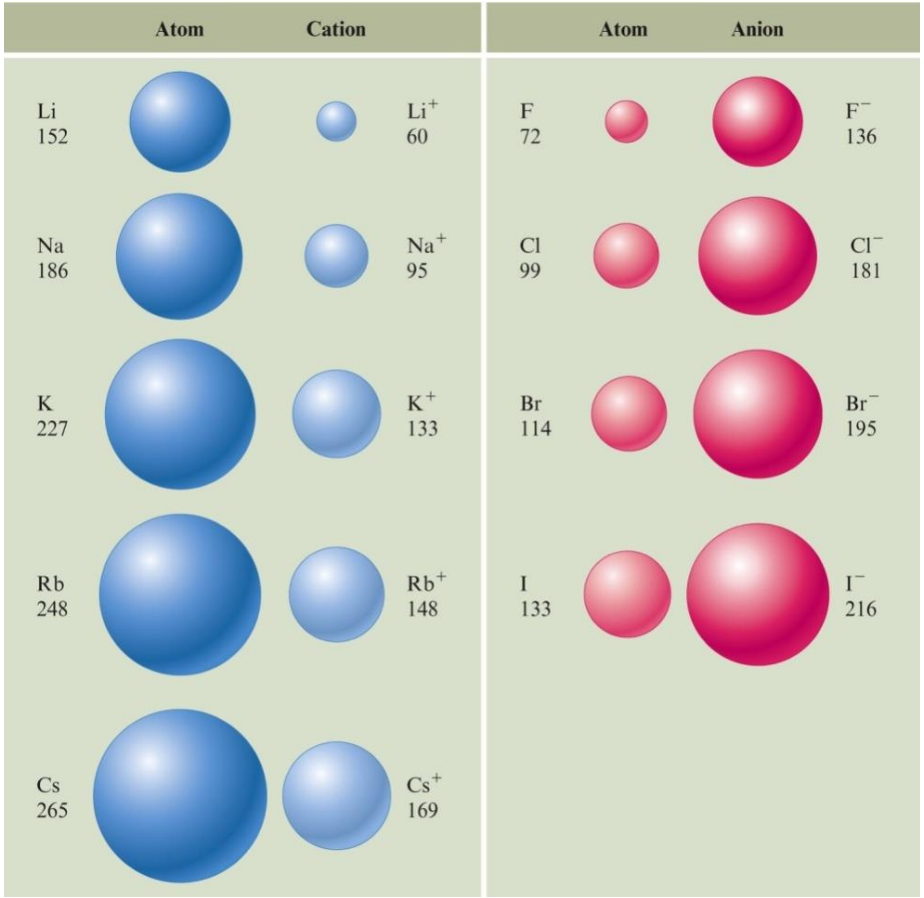
Structures of Ionic Compounds

- Ions are packed together to maximize the attractions between ions.



Characteristics of Ions and Ionic Compounds

Cations are always <u>smaller</u> than the parent atom.



Anions are always <u>larger</u> than the parent atom.

- Outer energy level was removed
- Unbalanced + charge draws e- closer to nucleus

•Excess negative charge weakens pull of nucleus on e- cloud

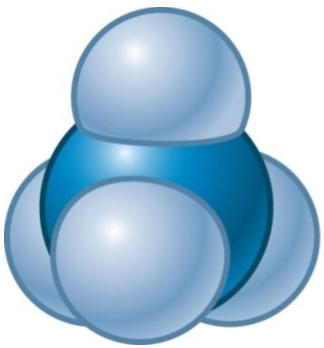
•Shielding by inner core e-

Characteristics of Ions and Ionic Compounds

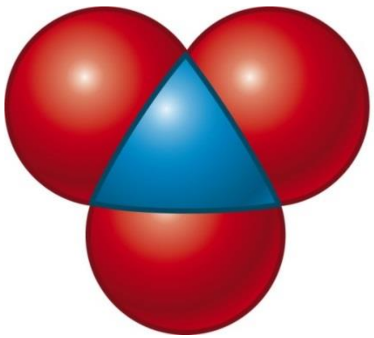
Ionic Compounds Containing Polyatomic Ions

- Polyatomic ions work in the same way as simple ions:

  The covalent bonds hold the polyatomic ion together so it behaves as a unit.



NH_4^+



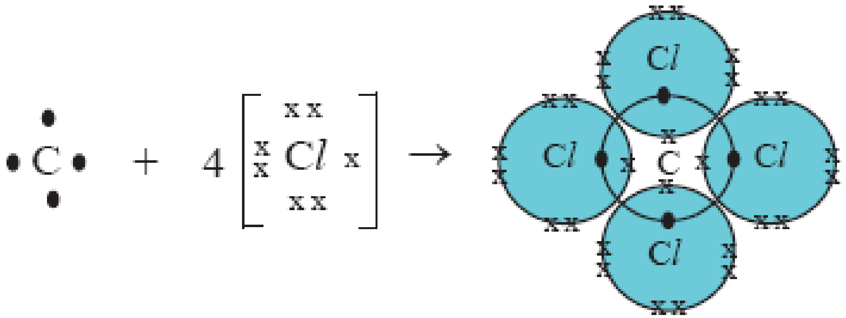
NO_3^-

Bonding

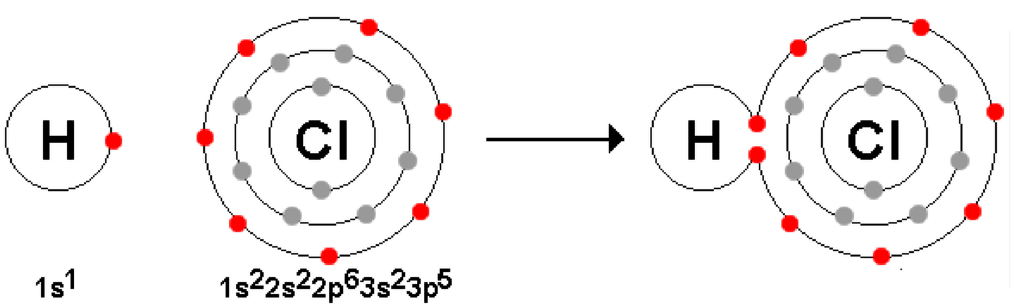
Covalent Bonding: results from atoms sharing electrons

Covalent compounds can form between

- **2 non-metals**



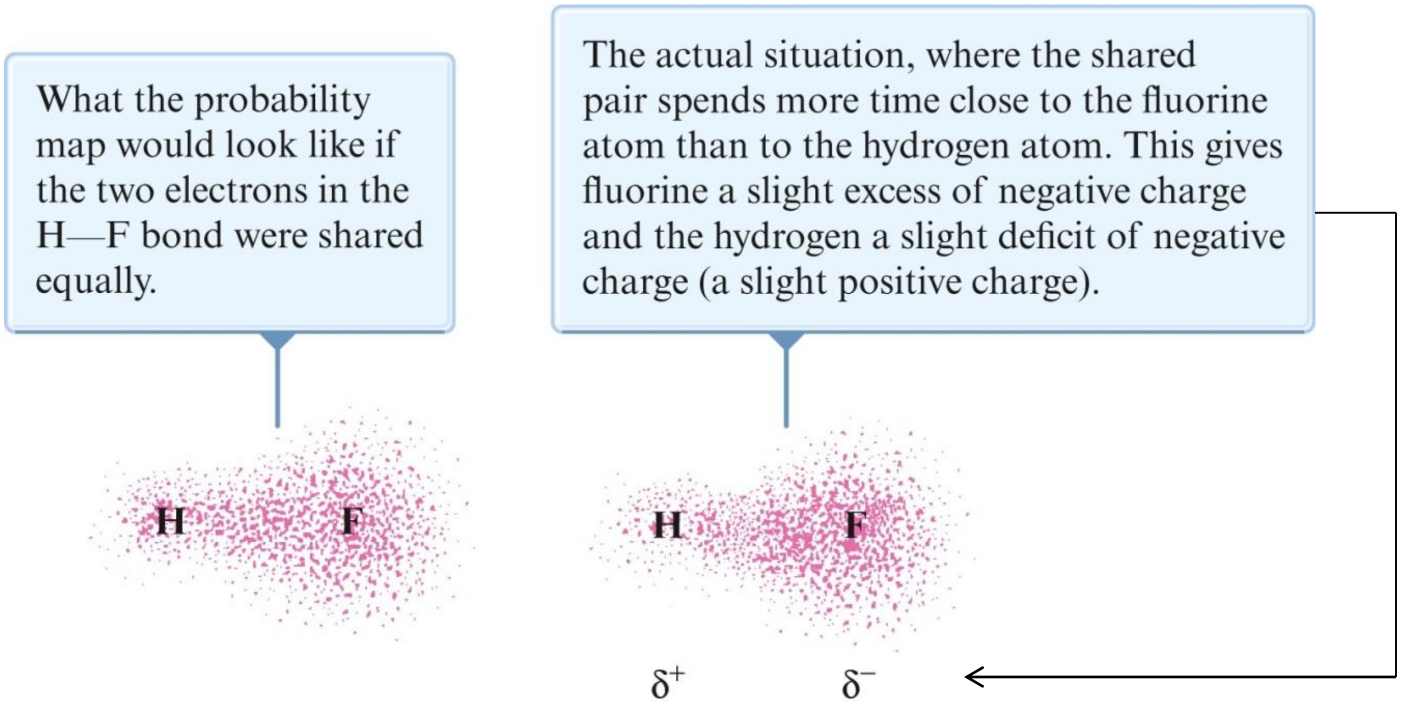
- **Hydrogen and a non-metal**



Bonding

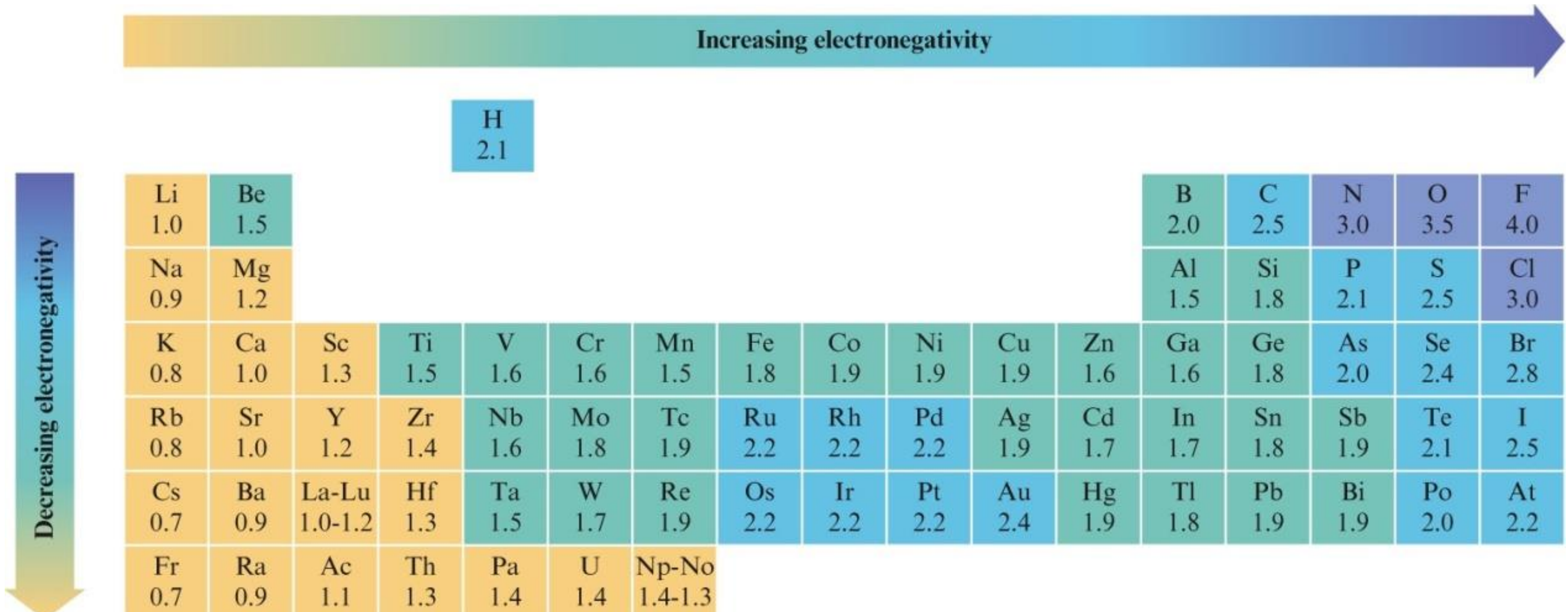
- A polar covalent bond results when electrons are shared unequally by nuclei.

    - One atom attracts the electrons more than the other atom.



Electronegativity

- **Electronegativity** – the relative ability of an atom in a molecule to attract shared electrons to itself (property of an atom in a molecule)
  - Increases from left to right across a period
  - Decreases down a group of representative elements



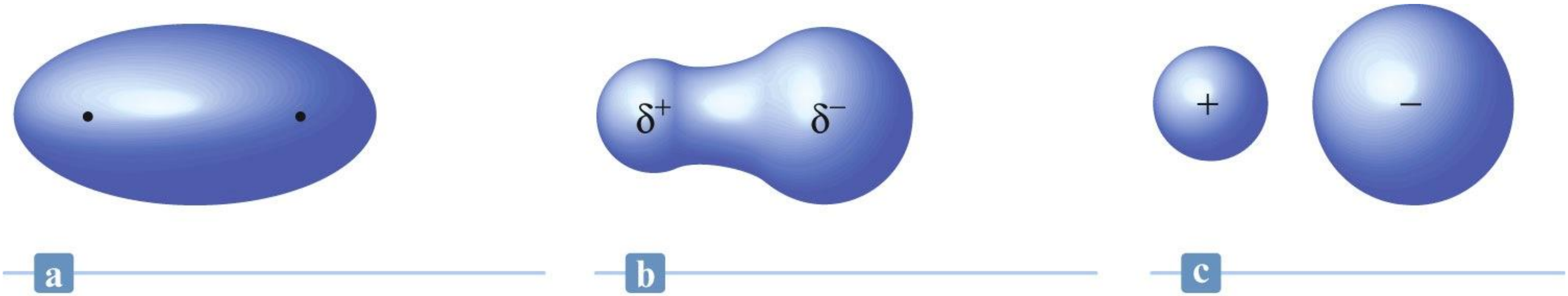
Electronegativity

The electronegativity difference between two atoms determines the type of bond which forms.

Type of Bond	Electronegativity Difference	Examples of Bonds
Nonpolar Covalent Bond	Less than 0.5	2.1 – 2.1 = 0 H-H
Polar Covalent Bond	0.5-1.9	3.5 – 2.5 = 1.0 C-O
Ionic Bond	2.0 and greater	3.0 - 0.9 = 2.1 Na-Cl

Electronegativity

- The polarity of a bond depends on the difference between the electronegativity values of the atoms forming the bond.



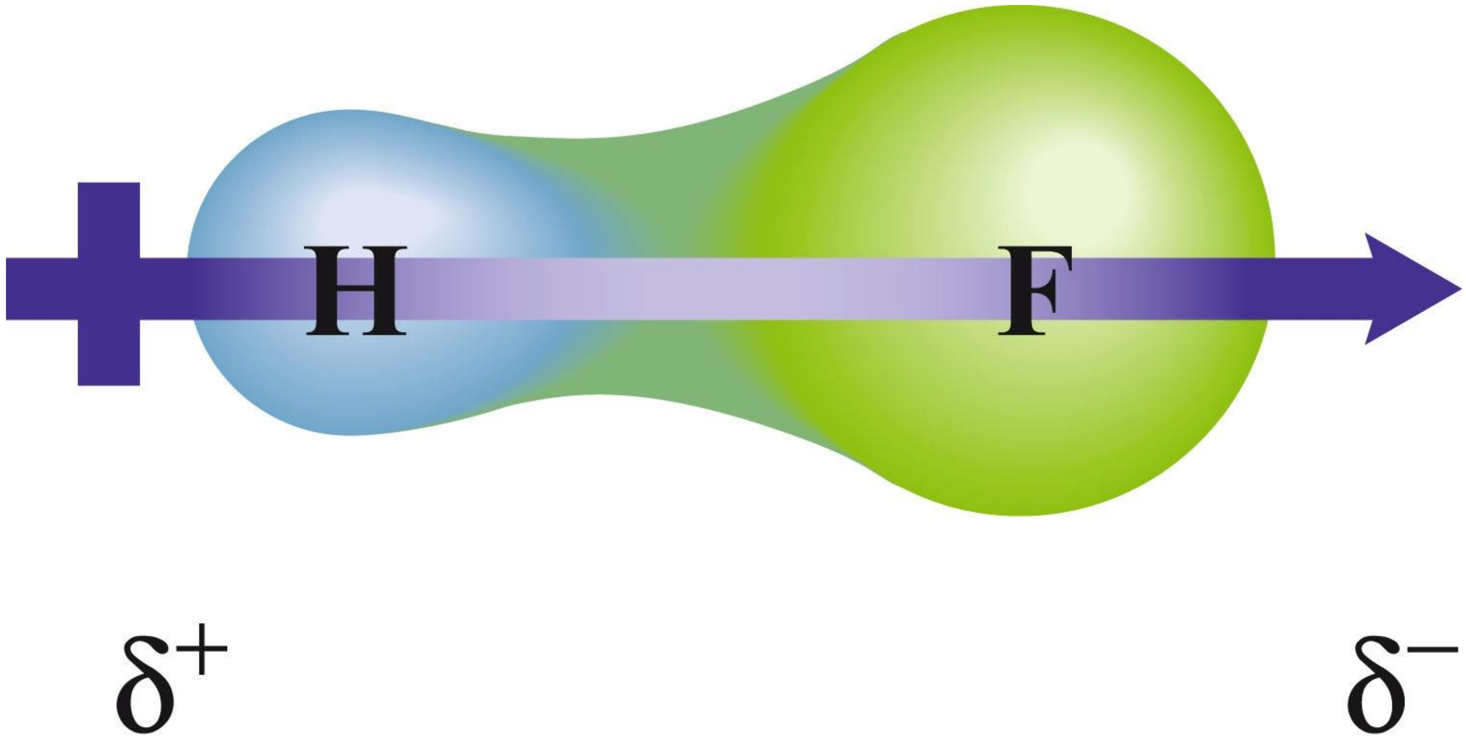
A covalent bond formed between identical atoms.

A polar covalent bond, with both ionic and covalent components.

An ionic bond, with no electron sharing.

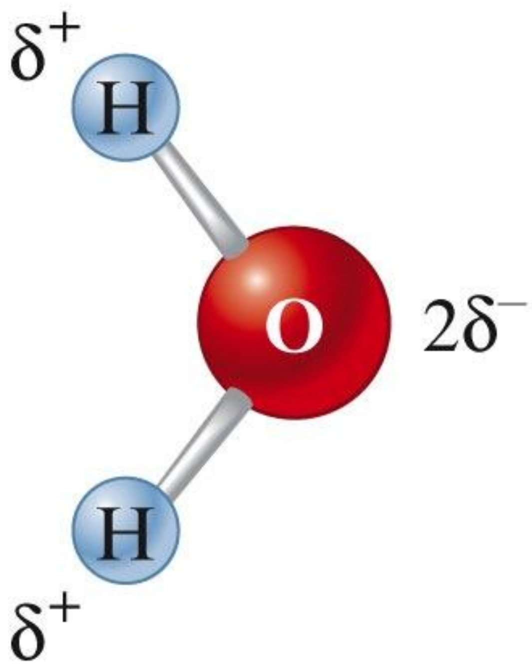
Bond Polarity and Dipole Moments

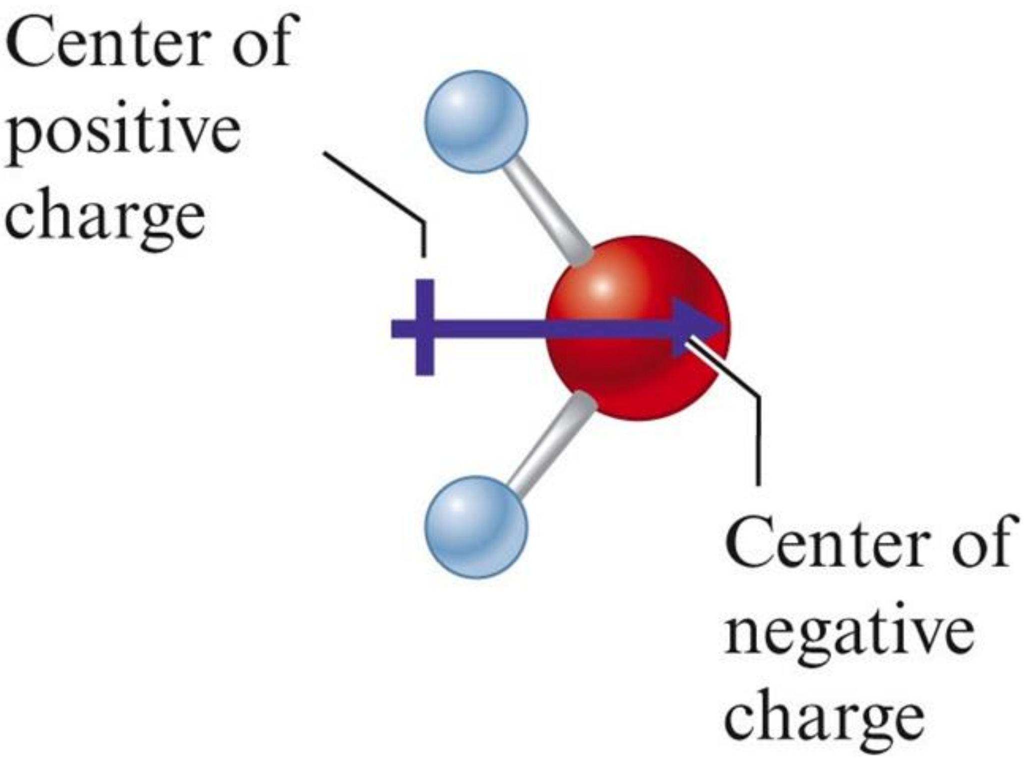
- A dipole moment results when a polar molecule has a center for positive charge separate from a center for negative charge.



Bond Polarity and Dipole Moments

Water molecule dipole moment:

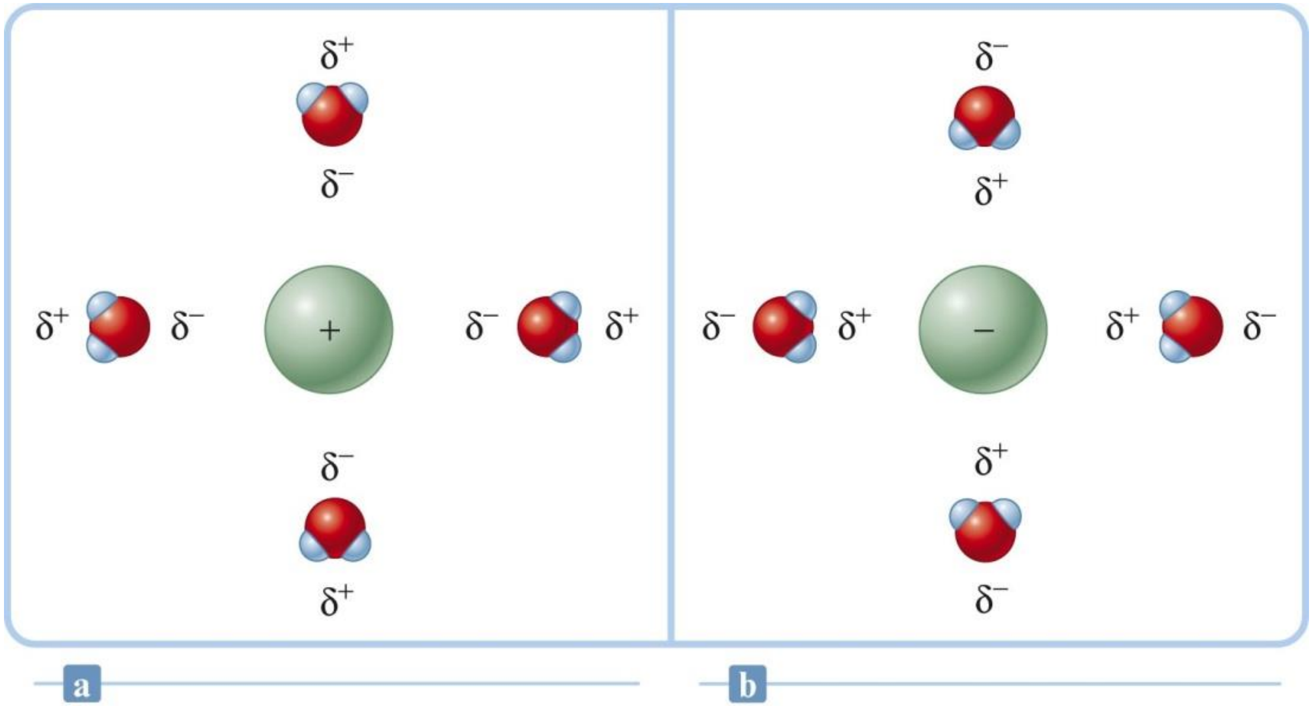




Bond Polarity and Dipole Moments

The polarity of water affects its properties:

- **Causes water to remain liquid at higher temperature**
- **Permits ionic compounds to dissolve in it**



Polar water molecules are strongly attracted to positive ions by their negative ends.

Water molecules are also strongly attracted to negative ions by their positive ends.

Lewis Structures

In covalent bonds the principle of achieving a noble gas configuration applies to the elements involved.

Duet Rule: The formation of a bond to fill the 1s shell (applies to Hydrogen)

Octet Rule: The second row non metals Carbon through Fluorine form molecules when their 2s and 2p (valence shells are filled.

Lewis Structures

Steps for writing Lewis structures:

1. Sum the valence electrons from all the atoms. Do not worry about keeping track of which electrons come from which atoms. The TOTAL number is what is important.

2. Use a pair of electrons to form a bond between each pair of bound atoms.

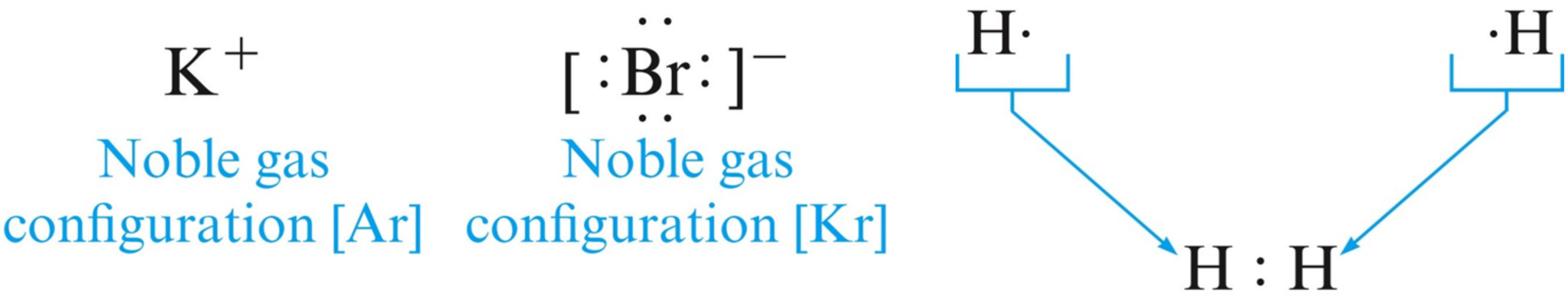
3. Arrange the remaining electrons to satisfy the Duet rule for hydrogen and the octet rule for the second row elements.

Lewis Structures

- Single bond – covalent bond in which 1 pair of electrons is shared by 2 atoms
- Double bond – covalent bond in which 2 pairs of electrons are shared by 2 atoms
- Triple bond – covalent bond in which 3 pairs of electrons are shared by 2 atoms

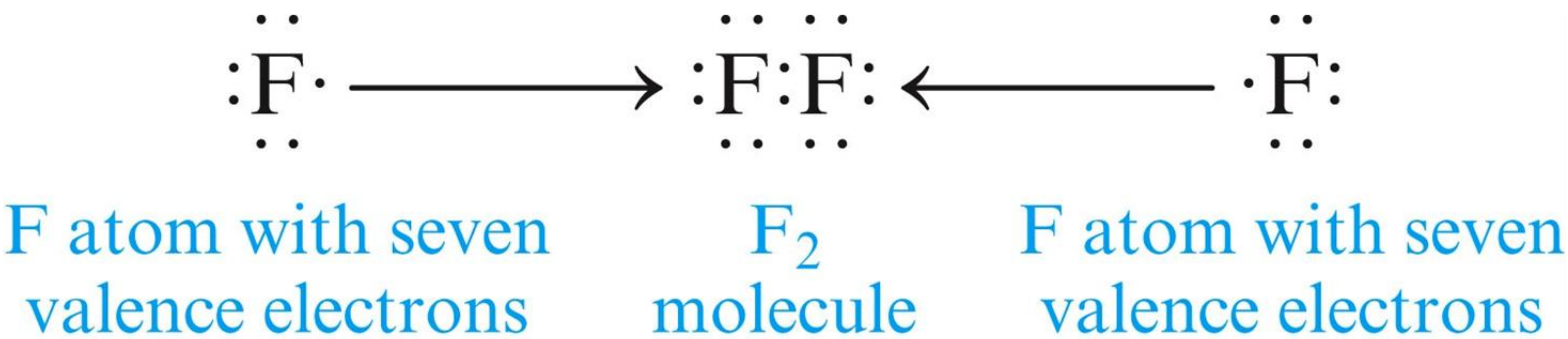
Lewis Structures

- In writing Lewis structures, we include only the valence electrons.
- Most important requirement
  - Atoms achieve noble gas electron configuration (octet rule, duet rule).

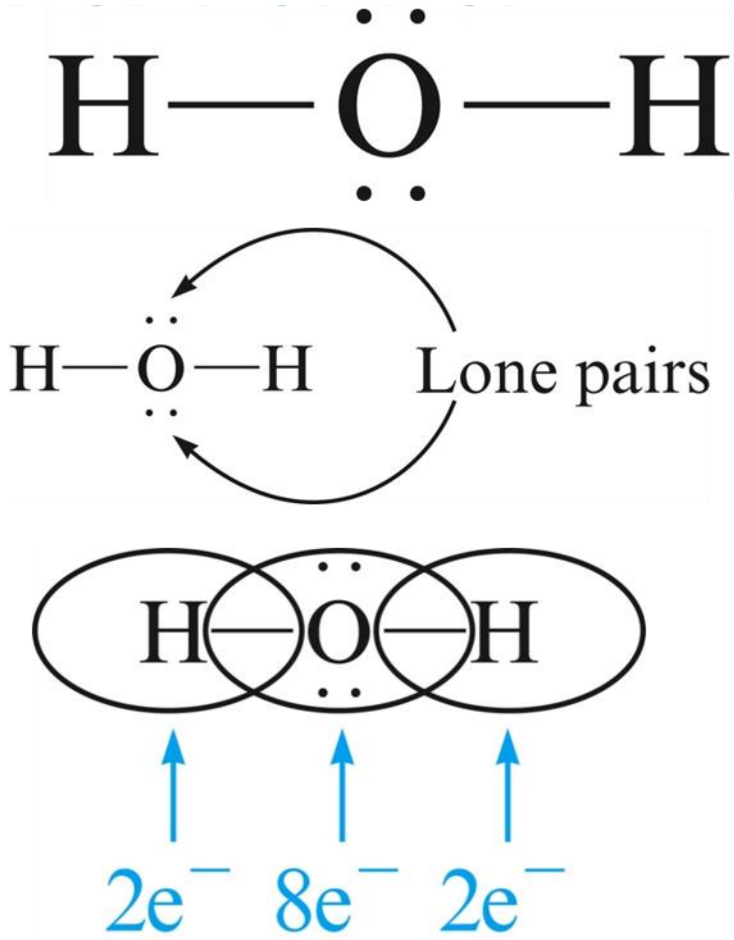


Lewis Structures

- Bonding pairs are shared between 2 atoms.
- Unshared pairs (lone pairs) are not shared and not involved in bonding.

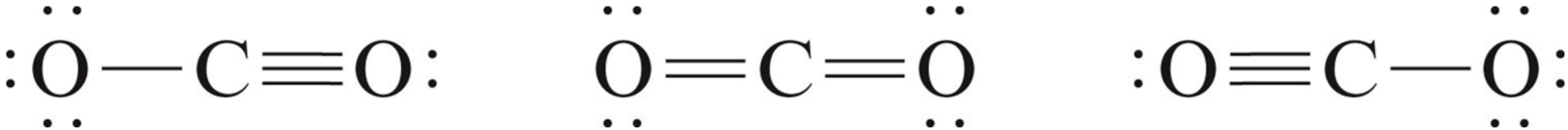


Lewis Structures



Lewis Structures

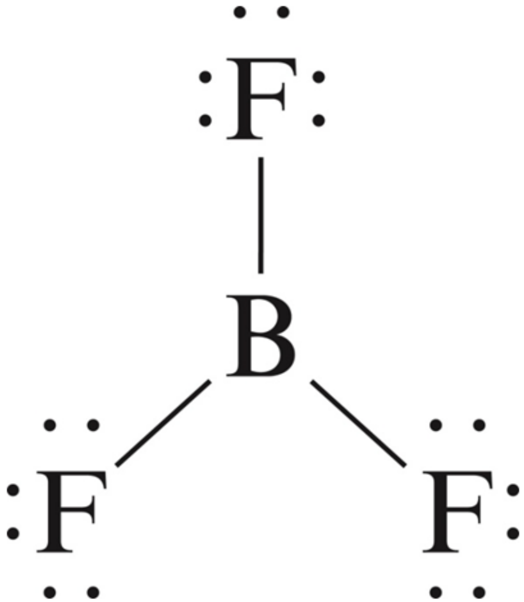
- A molecule shows resonance when more than one Lewis structure can be drawn for the molecule.



Lewis Structures

Some Exceptions to the Octet Rule

- Boron – incomplete octet



- Molecules containing odd numbers of electrons – NO and NO_2

Formal Charge

- Can be used to distinguish between several possible Lewis structures.

- Generally, the Lewis structure with the smallest formal charges on individual atoms will be the correct one.

- Determine electron distribution within a molecule or ion, and therefore give us a starting point to predict chemical and physical properties

- Charge distribution (identification of atoms that are electron rich or electron poor) can be useful to understand how and why reactions occur or how molecules interact with each other

Formal Charge

<u>Formal Charge</u>:

Formal Charge = (number of valence electrons on free atom) – (number of lone pair electrons) - (shared electrons/2)

where FC =formal charge

$$FC = V - N - B/2$$

V = valence electrons

N= lone pair electrons

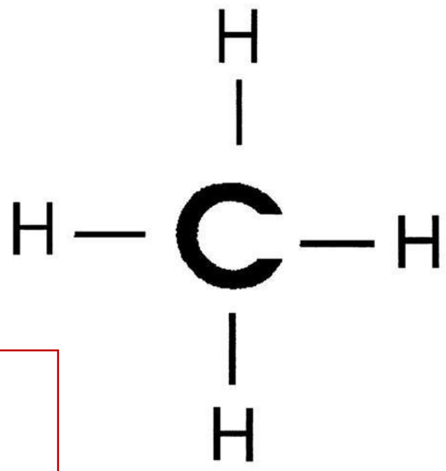
B = shared electrons (in bonds)

Assumptions:

- Lone pair electrons belong entirely to the atom in question
- Shared electrons are divided equally between the two sharing atoms

Formal Charge

Carbon in <u>methane</u>:

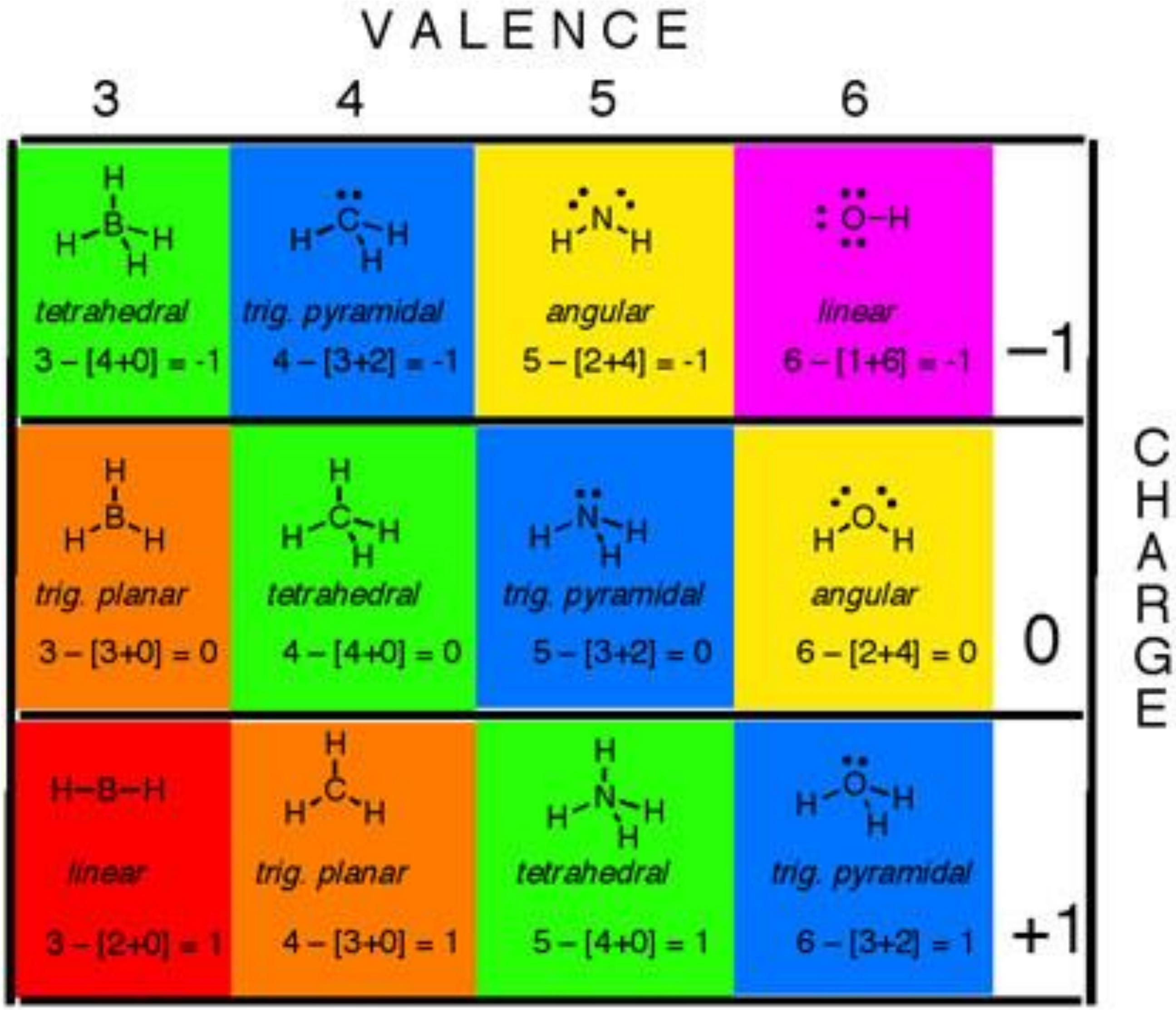


Valence electrons (V)	= 4
Lone pair electrons (N)	= 0
Shared electrons (B)	= 8

FC = 4 − 0 − (8÷2) = 0

Formal charge = 0

Formal Charge



Structures of Molecules

B. The VSEPR Model

- Valence shell electron pair repulsion (VSEPR) model
  - Molecular structure is determined by minimizing repulsions between electron pairs.

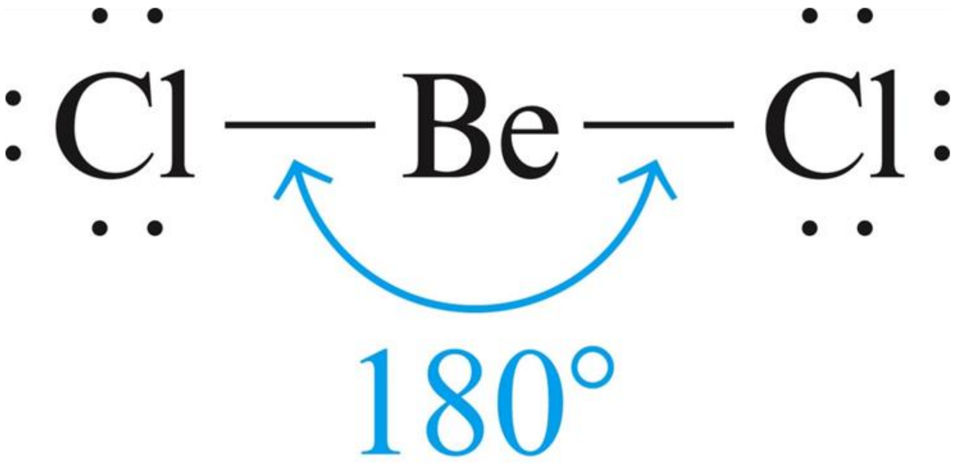
Predicting structures of molecules

1. Draw the Lewis Structure for the molecule.
2. Count the electron pairs and arrange them in the way that minimize repulsions (put the pairs as far apart as possible).
3. Determine the positions of the atoms from the way the electrons pairs are shared.
4. Determine the name of the molecular structure from the positions of the atoms.
5. Must Show 3-D shape by using  and 

Structures of Molecules

Two Pairs of Electrons

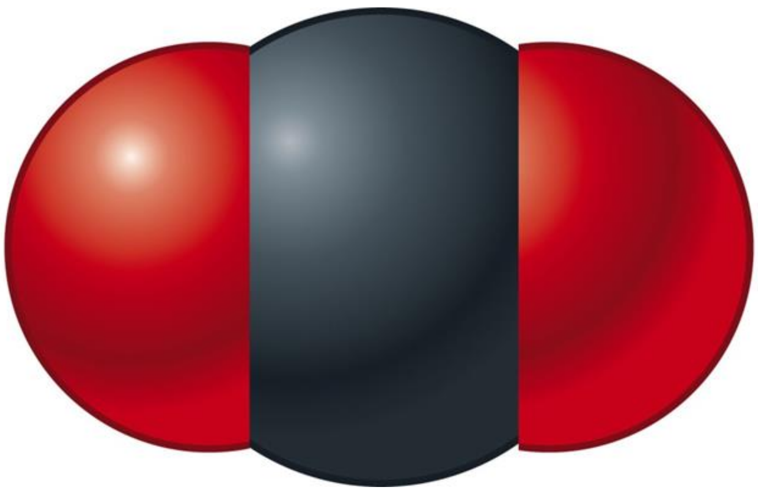
- $BeCl_2$

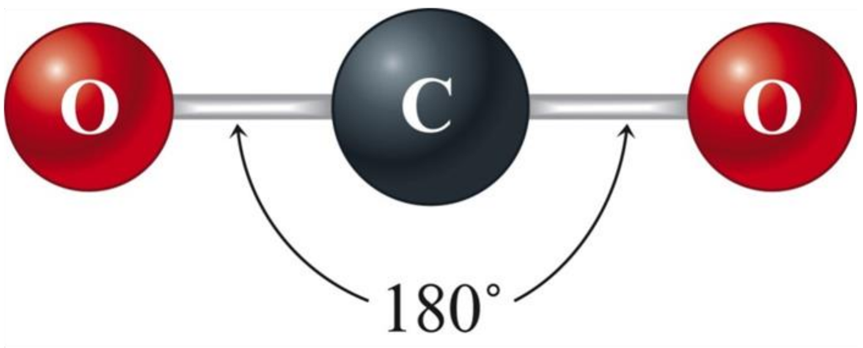


  - 180^o - linear

Structures of Molecules

- Linear structure – atoms in a line
  - Carbon dioxide

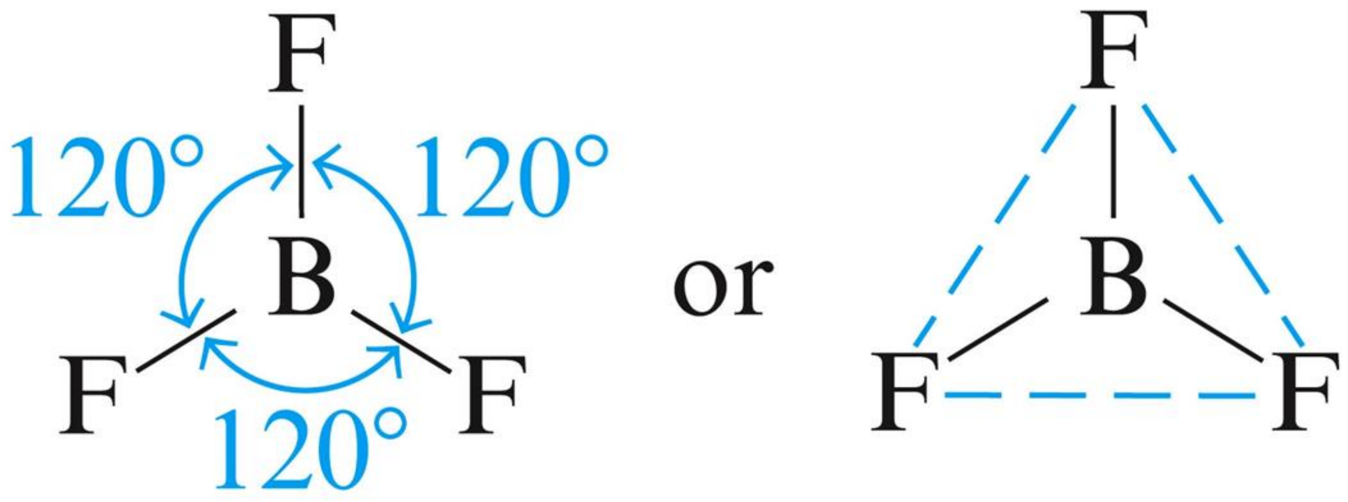




Structures of Molecules

Three Pairs of Electrons

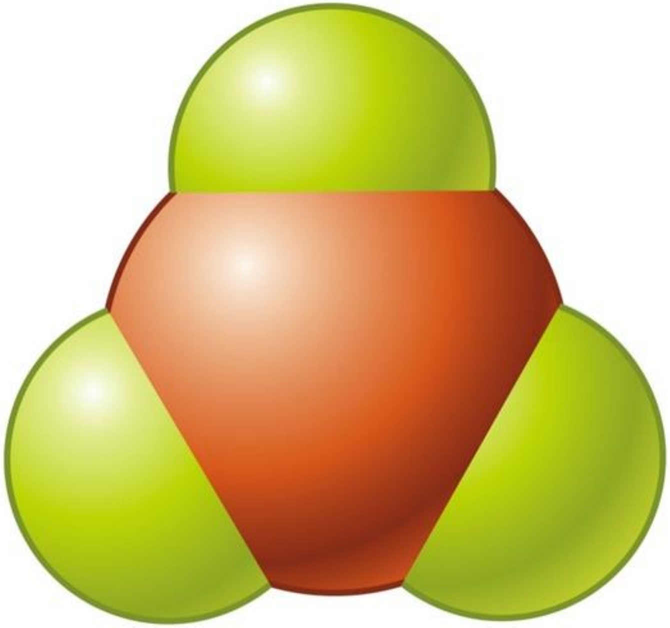
- BF_3

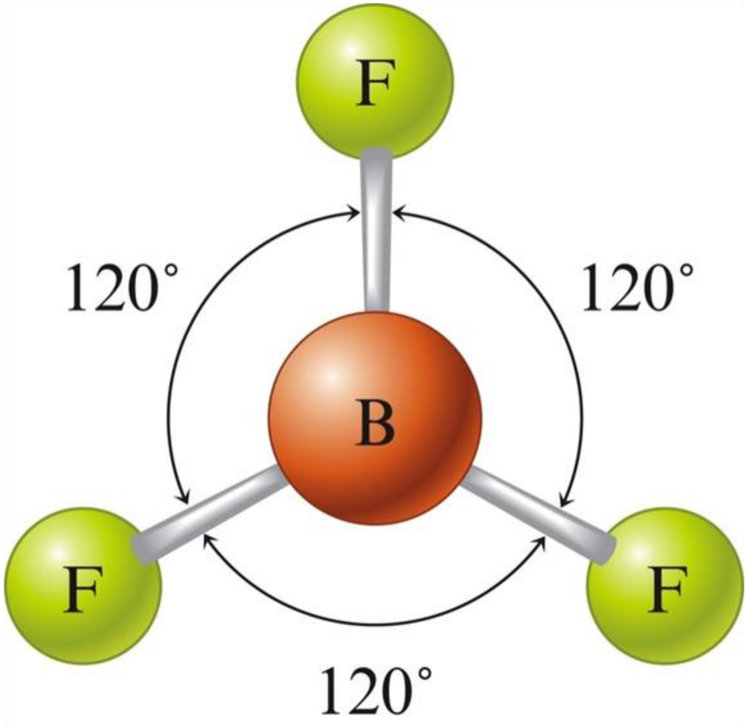


  - 120º – trigonal planar

Structures of Molecules

- Trigonal planar – atoms in a triangle
  - BF_3

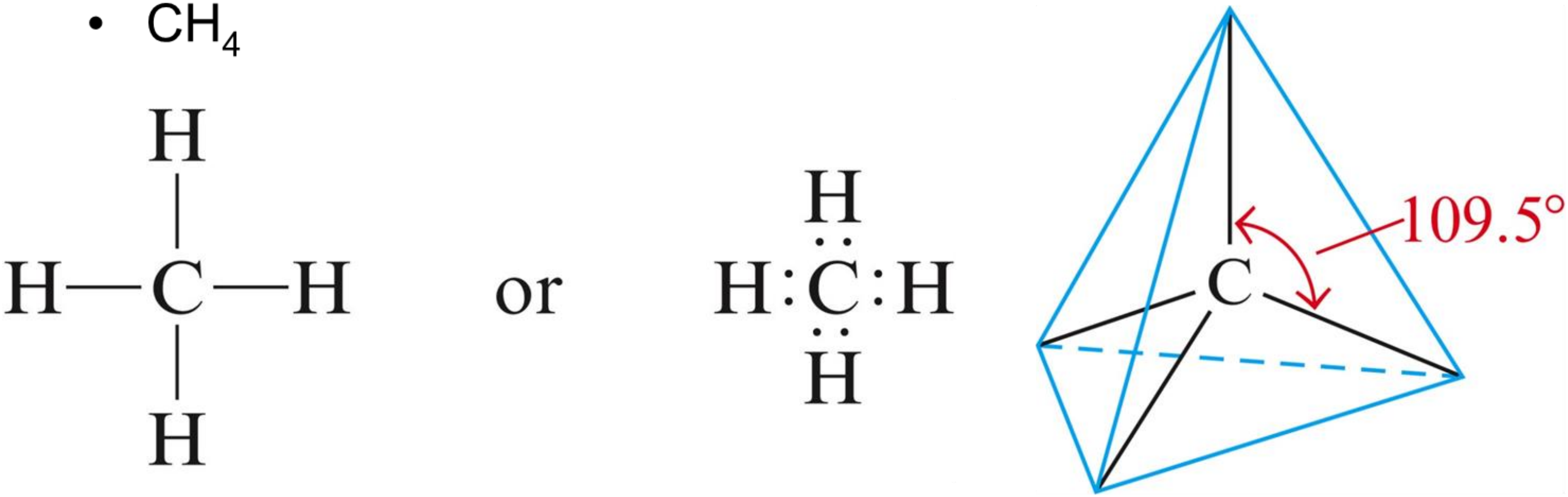




Structures of Molecules

Four Pairs of Electrons

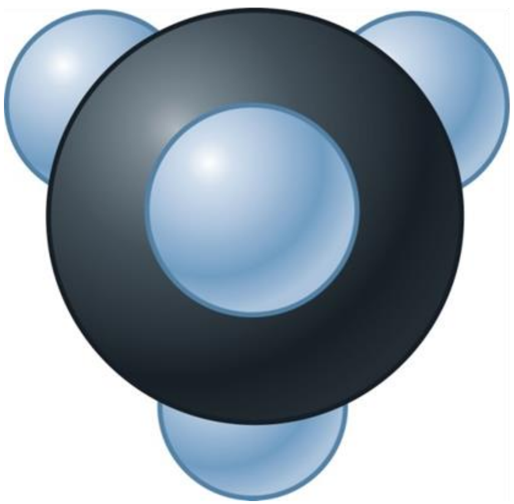
- CH_4

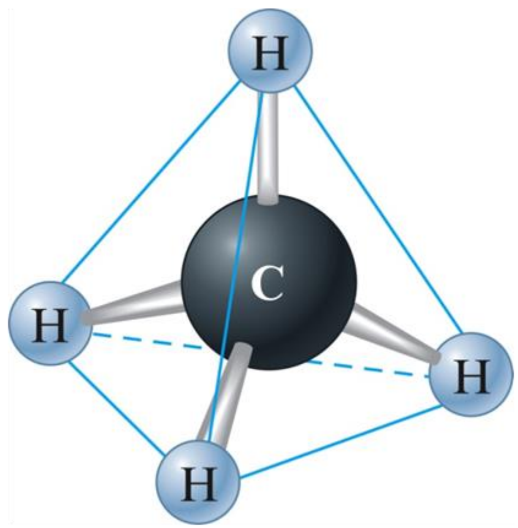


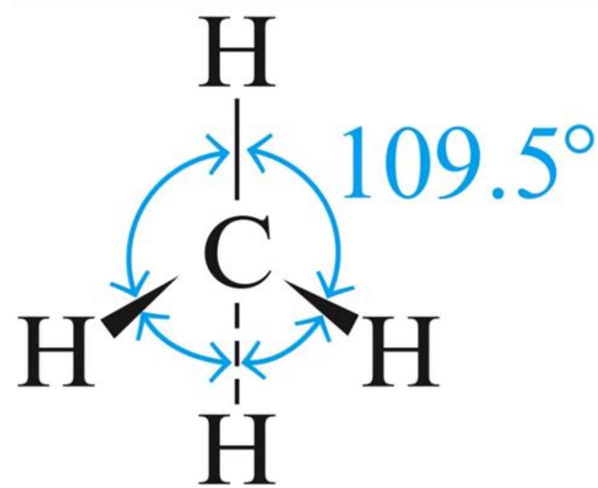
  - 109.5° – tetrahedral

Structures of Molecules

- Tetrahedral structure
  - methane

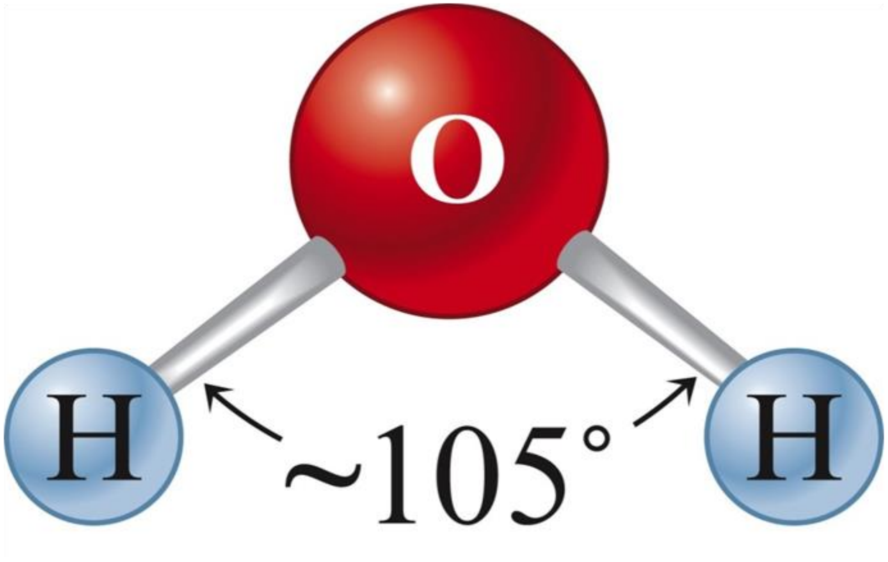






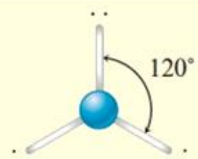
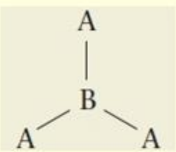
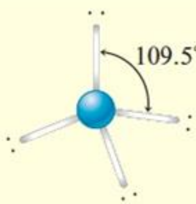
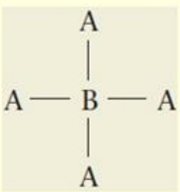
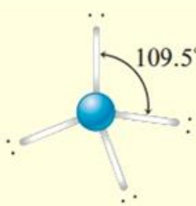
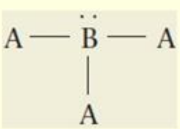
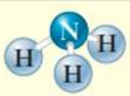
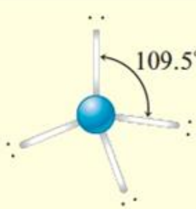
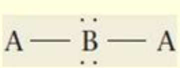
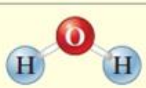
Structures of Molecules

- Three dimensional arrangement of the atoms in a molecule
  - Water - bent

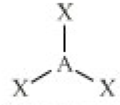


Structures of Molecules

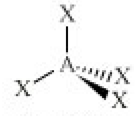
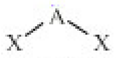
Table 12.4 Arrangements of Electron Pairs and the Resulting Molecular Structures for Two, Three, and Four Electron Pairs

Number of Electron Pairs	Bonds	Electron Pair Arrangement	Ball-and-Stick Model	Molecular Structure	Partial Lewis Structure	Ball-and-Stick Model
2	2	Linear		Linear	A—B—A	
3	3	Trigonal planar (triangular)		Trigonal planar (triangular)		
4	4	Tetrahedral		Tetrahedral		
4	3	Tetrahedral		Trigonal pyramid		
4	2	Tetrahedral		Bent or V-shaped		

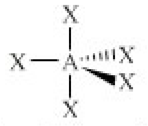
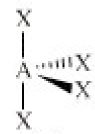
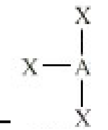
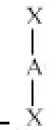
VSEPR

Formula	Lone Pairs	Structure	Shape
BA	0		Linear
BA_2	0		Linear
BAE	1		Linear
BA_3	0		Trigonal Planar
BA_2E	1		Bent
BAE_2	2		Linear

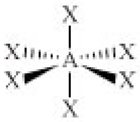
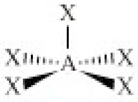
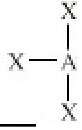
VSEPR

Formula	Lone Pairs	Structure	Shape
BA_4	0		Tetrahedral
BA_3E	1		Pyramidal
BA_2E_2	2		Bent
BAE_3	3		Linear

VSEPR

Formula	Lone Pairs	Structure	Shape
BA_5	0		Trigonal Bipyramidal
BA_4E	1		Seesaw
BA_3E_2	2		T-Shaped
BA_2E_3	3		Linear
BAE_4	4		Linear

VSEPR

Formula	Lone Pairs	Structure	Shape
BA_6	0		Octahedral
BA_5E	1		Square Pyramidal
BA_4E_2	2		Square Planar
BA_3E_3	3		T-Shape
BA_2E_4	4		Linear
BAE_5	5		Linear