Zumdahl • Zumdahl • DeCoste

World of CHEMISTRY

Chapter 12

Chemical Bonding

Chapter 12 Overview

- Ionic & covalent bonds & their formation
- Polar covalent bonds
- Bond relationships to electronegativity
- Bond polarity & its relationship to molecular polarity
- Ionic structures
- Ionic size
- Lewis structures
- Molecular structures & bond angles
- VSEPR model

Why are graphite and diamonds so different even though they are both carbon?

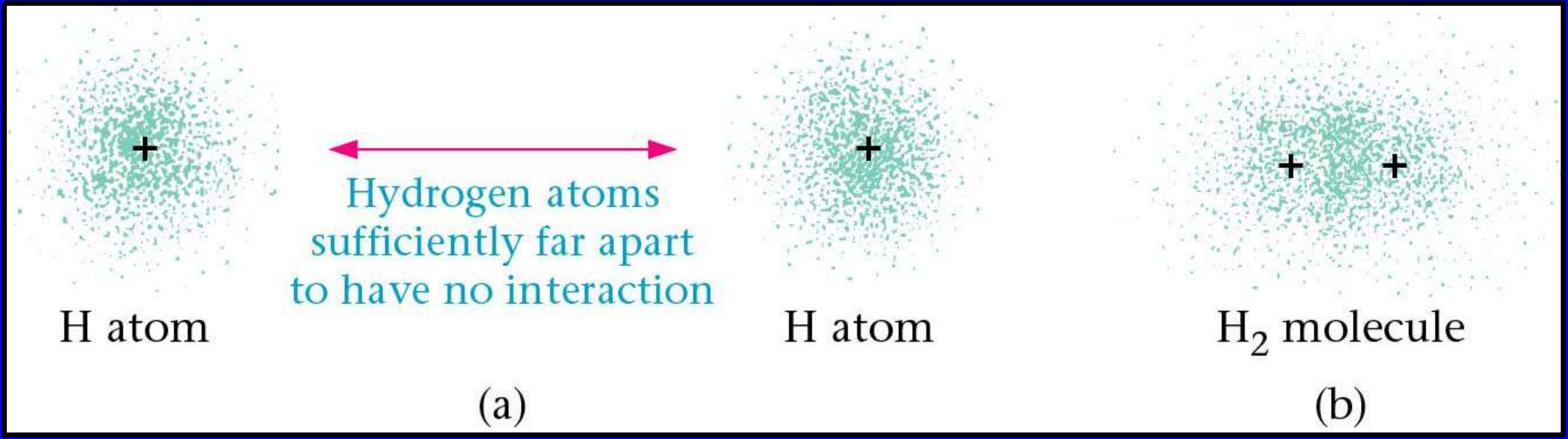
- The carbon atoms are bound differently

- Structure & shape very important
  - Different properties
  - Reactions
  - Smell
  - Taste

Types of Chemical Bonds

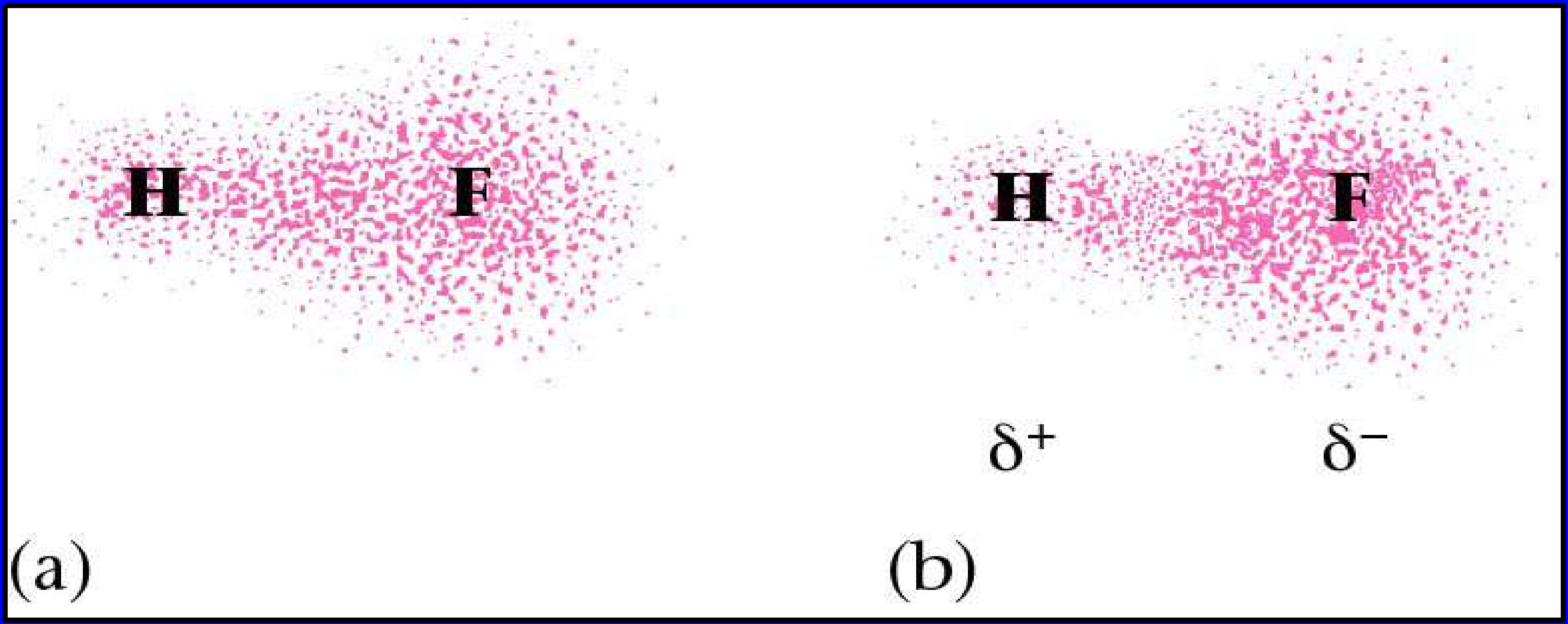
- ***Bond:*** force that holds groups of 2 or more atoms together and makes them function as a unit
- ***Bond Energy:*** the energy required to break a bond
- ***Ionic Bonding:*** attraction between oppositely charged ions
- ***Ionic Compound:*** result of metal reacting with a nonmetal
- ***Covalent Bonding***: bonding where atoms share electrons
- ***Polar Covalent Bond:*** covalent bond where electrons are not shared equally because on atom attracts them more strongly than the other

Figure 12.1: The formation of a bond between two atoms.



Covalent Bond

Figure 12.2: Probability representations of the electron sharing in HF.



Polar Covalent Bond

Electronegativity

- *The relative ability of an atom in a molecule to attract shared electrons to itself*
- Chemists determine by measuring the polarities of bonds between different atoms
- Higher value = higher attraction of electrons
- Difference between electronegativity values determines polarity of molecule
  - Larger difference = more polar
  - Smaller difference = more equally shared
  - If difference is greater than 2 – bond is ionic (electrons are transferred)

Table 12.1

TABLE 12.1

The Relationship Between Electronegativity and Bond Type

Electronegativity Difference Between the Bonding Atoms	Bond Type	Covalent Character	Ionic Character
Zero	Covalent		
↓	↓	↑	↓
Intermediate	Polar covalent	Increase	Increases
↓	↓	↑	↓
Large	Ionic		

Figure 12.3: Electronegativity values for selected elements.

Increasing electronegativity

Decreasing electronegativity

				H 2.1												
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La-Lu 1.0-1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac 1.1	Th 1.3	Pa 1.4	U 1.4	Np-No 1.4-1.3										

Key

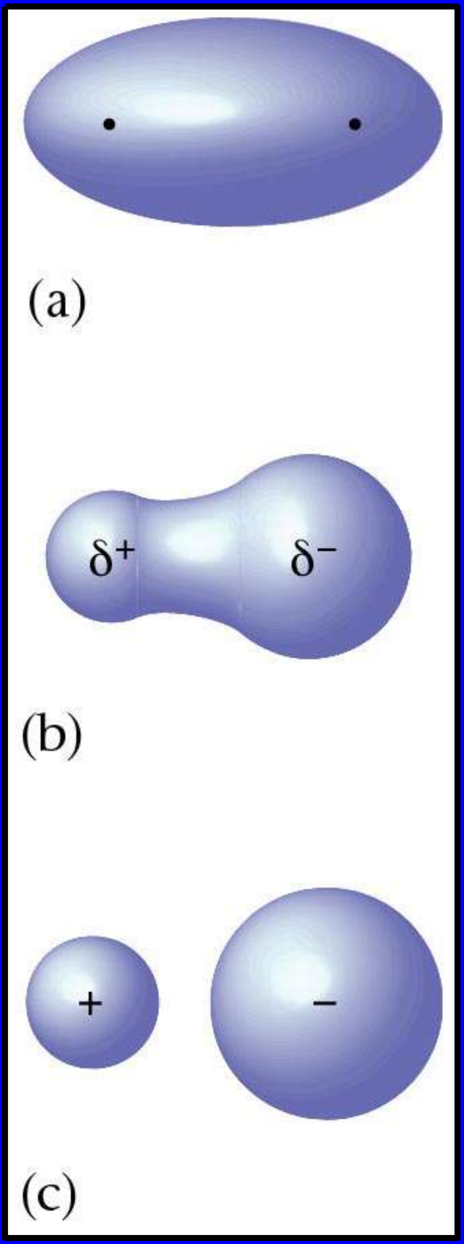
- < 1.5
- 1.5–1.9
- 2.0–2.9
- 3.0–4.0

Figure 12.4:
The three possible types of bonds.

(a) Covalent

(b) Polar covalent

(c) Ionic



Bond Polarity & Dipole Moments

- ***Dipole moment:*** property of a molecule whereby the charge distribution can be represented by a center of positive charge and a center of negative charge
  - Represent with arrow pointing toward center of negative charge
  - All diatomic polar molecules have a dipole moment
- Dipole moments strongly effect properties
  - Polarity of water crucial to life
  - Allow materials to dissolve in water (attract + & -)
  - Water molecules attracted to each other (higher boiling point – keeps water in on Earth from evaporating)

Figure 12.5: Charge distribution in the water molecule.

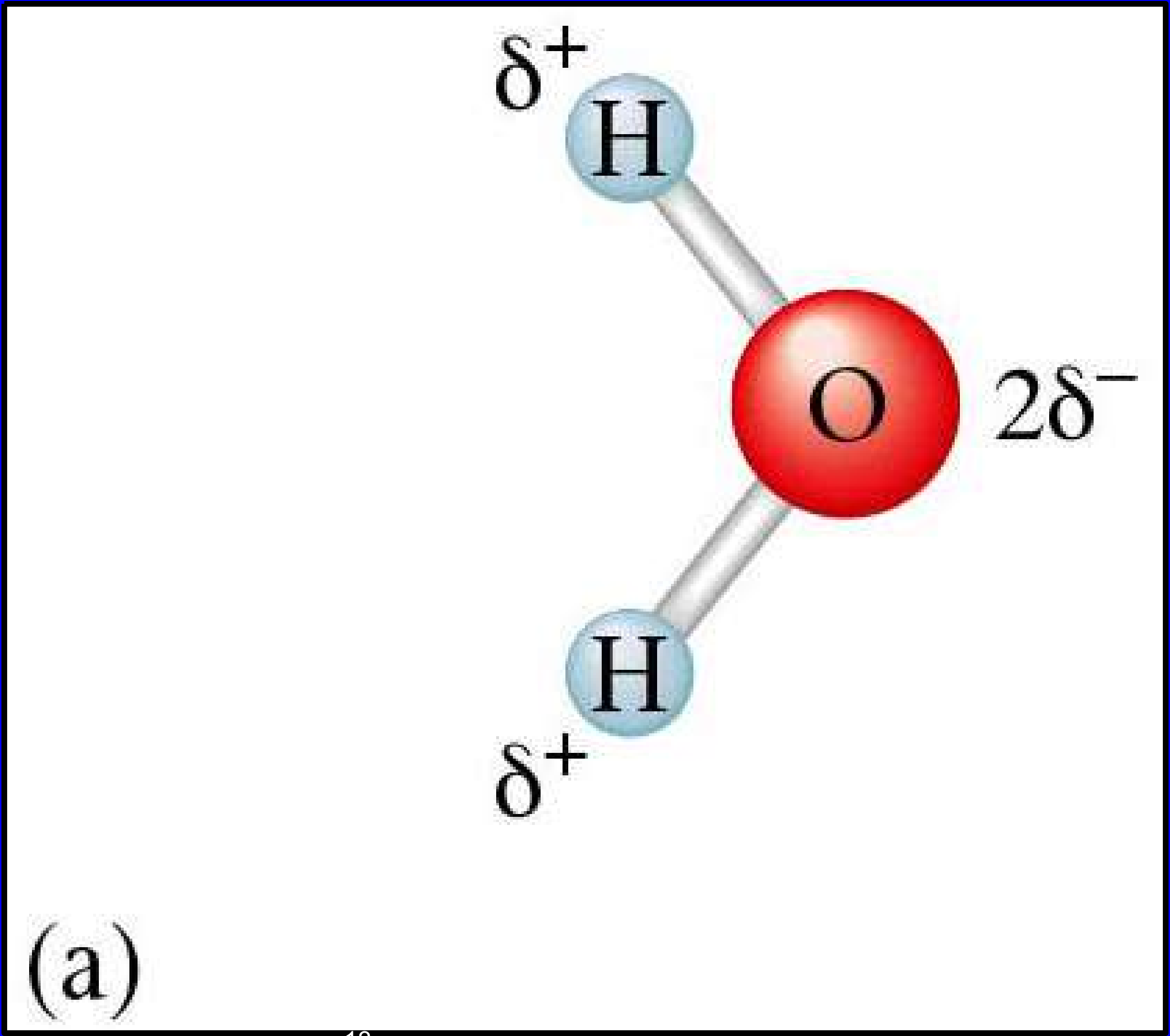


Figure 12.5: Water molecule behaves as if it had a positive and negative end.

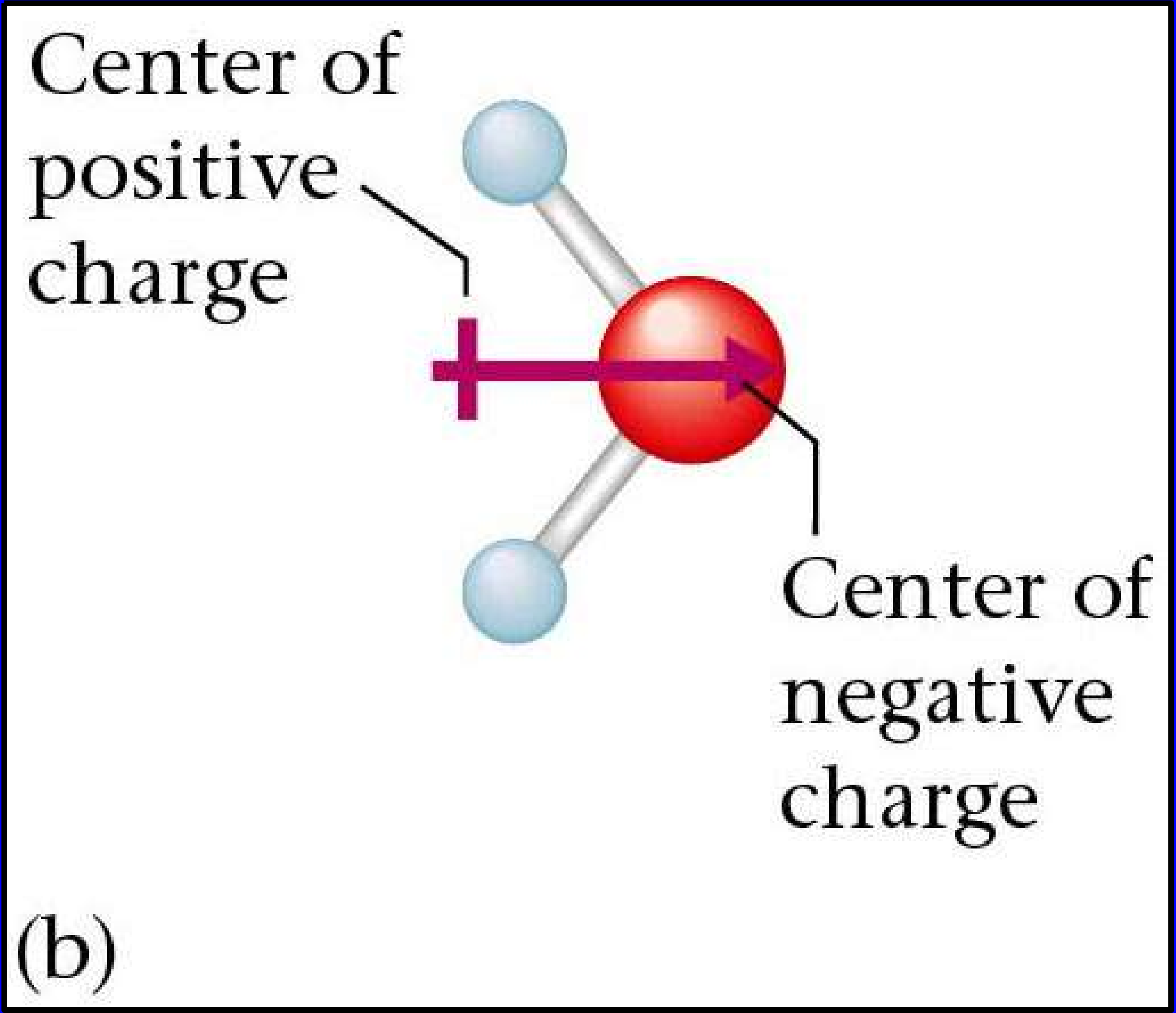


Figure 12.6:

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

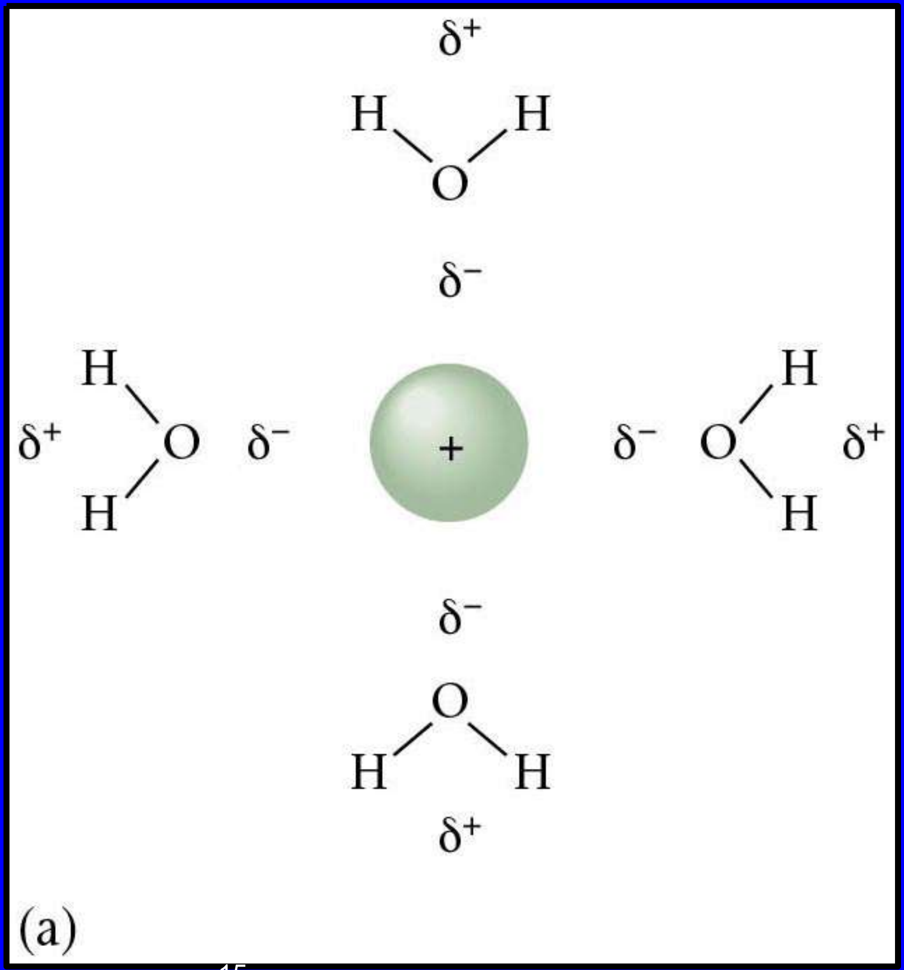
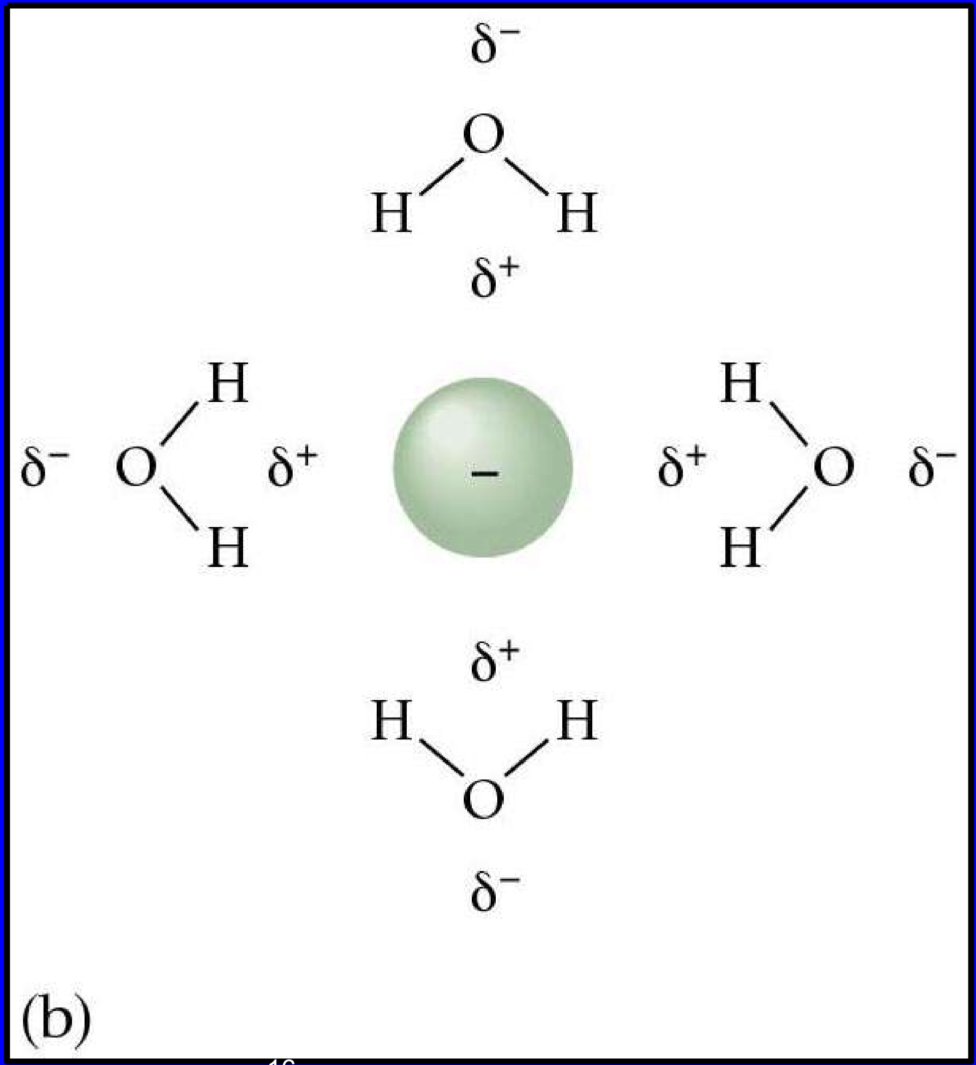


Figure 12.6: Polar water molecules are strongly attracted to negative ions by their positive ends.



Electron Configurations of Ions

- Main group metals form ions by losing enough electrons to achieve the configuration of the previous noble gas (transition metals behavior is more complicated)
- Nonmetals form ions by gaining enough electrons to achieve the configuration of the next noble gas

Table 12.2

TABLE 12.2

The Formation of Ions by Metals and Nonmetals

Group	Ion Formation	Electron Configuration: Atom		Electron Configuration: Ion
1	$Na \rightarrow Na^{+} + e^{-}$	$[Ne]3s^1$	$\xrightarrow{e^{-}\text{ lost}}$	[Ne]
2	$Mg \rightarrow Mg^{2+} + 2e^{-}$	$[Ne]3s^2$	$\xrightarrow{2e^{-}\text{ lost}}$	[Ne]
3	$Al \rightarrow Al^{3+} + 3e^{-}$	$[Ne]3s^23p^1$	$\xrightarrow{3e^{-}\text{ lost}}$	[Ne]
6	$O + 2e^{-} \rightarrow O^{2-}$	$[He]2s^22p^4 + 2e^{-}$	$\rightarrow [He]2s^22p^6 =$	[Ne]
7	$F + e^{-} \rightarrow F^{-}$	$[He]2s^22p^5 + e^{-}$	$\rightarrow [He]2s^22p^6 =$	[Ne]

Table 12.3

TABLE 12.3

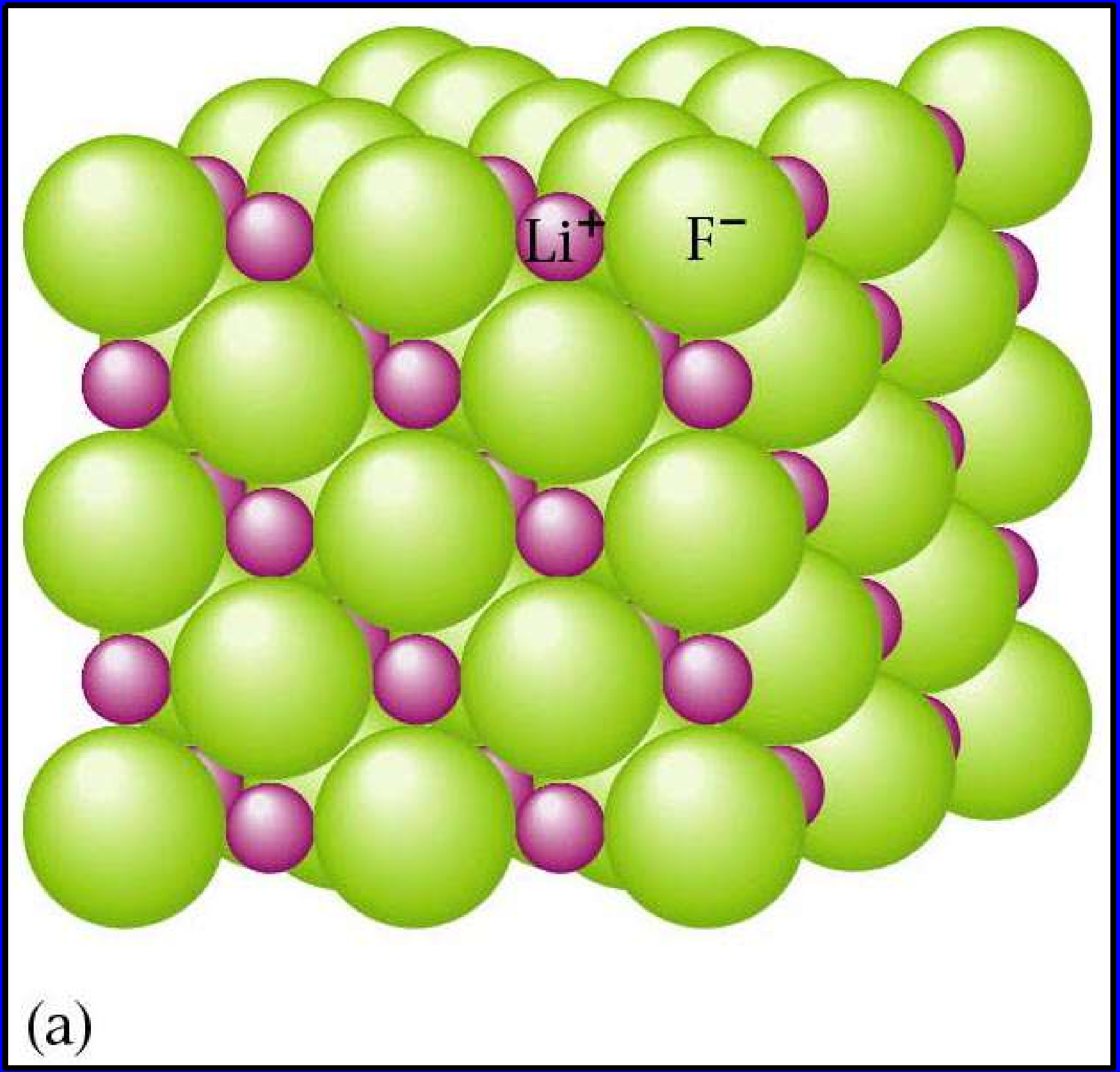
Common Ions with Noble Gas Configurations in Ionic Compounds

Group 1	Group 2	Group 3	Group 6	Group 7	Electron Configuration
Li^{+}	Be^{2+}				[He]
Na^{+}	Mg^{2+}	Al^{3+}	O^{2-}	F^{-}	[Ne]
K^{+}	Ca^{2+}		S^{2-}	Cl^{-}	[Ar]
Rb^{+}	Sr^{2+}		Se^{2-}	Br^{-}	[Kr]
Cs^{+}	Ba^{2+}		Te^{2-}	I^{-}	[Xe]

Electron Configurations & Bonding

- Ionic compounds react to form binary compounds with the ions having electron configurations of noble gases
- When two nonmetals react to form a covalent bond, they share electrons in a way that completes the valence-electron configurations of both atoms (they attain noble gas configurations).

Figure 12.8: Ions as packed spheres.



LiF is empirical or simplest formula for Lithium Fluoride

Actual compound contains huge & = numbers of Li^+ and F^- ions packed together

Figure 12.8: Positions (centers) of the ions.

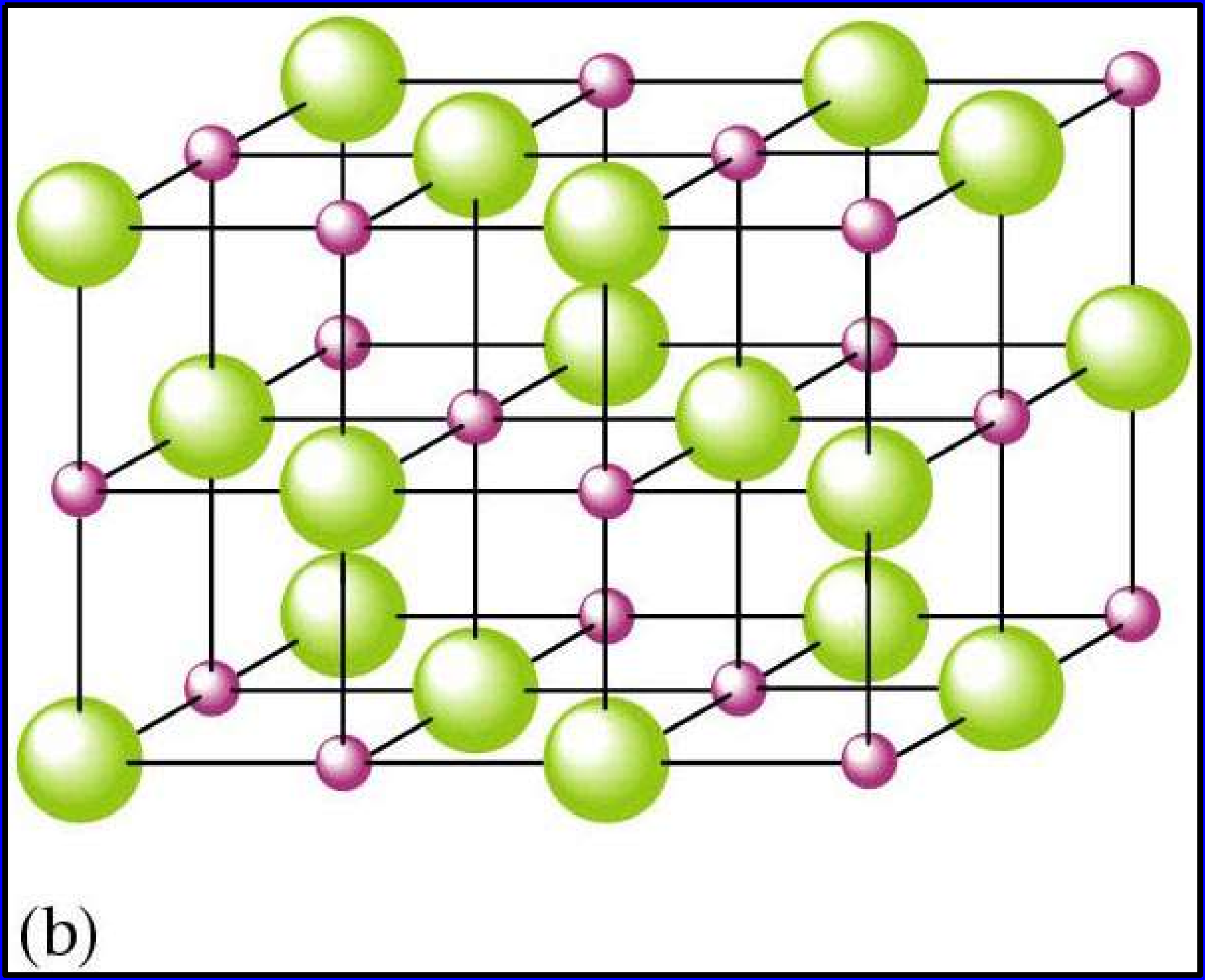
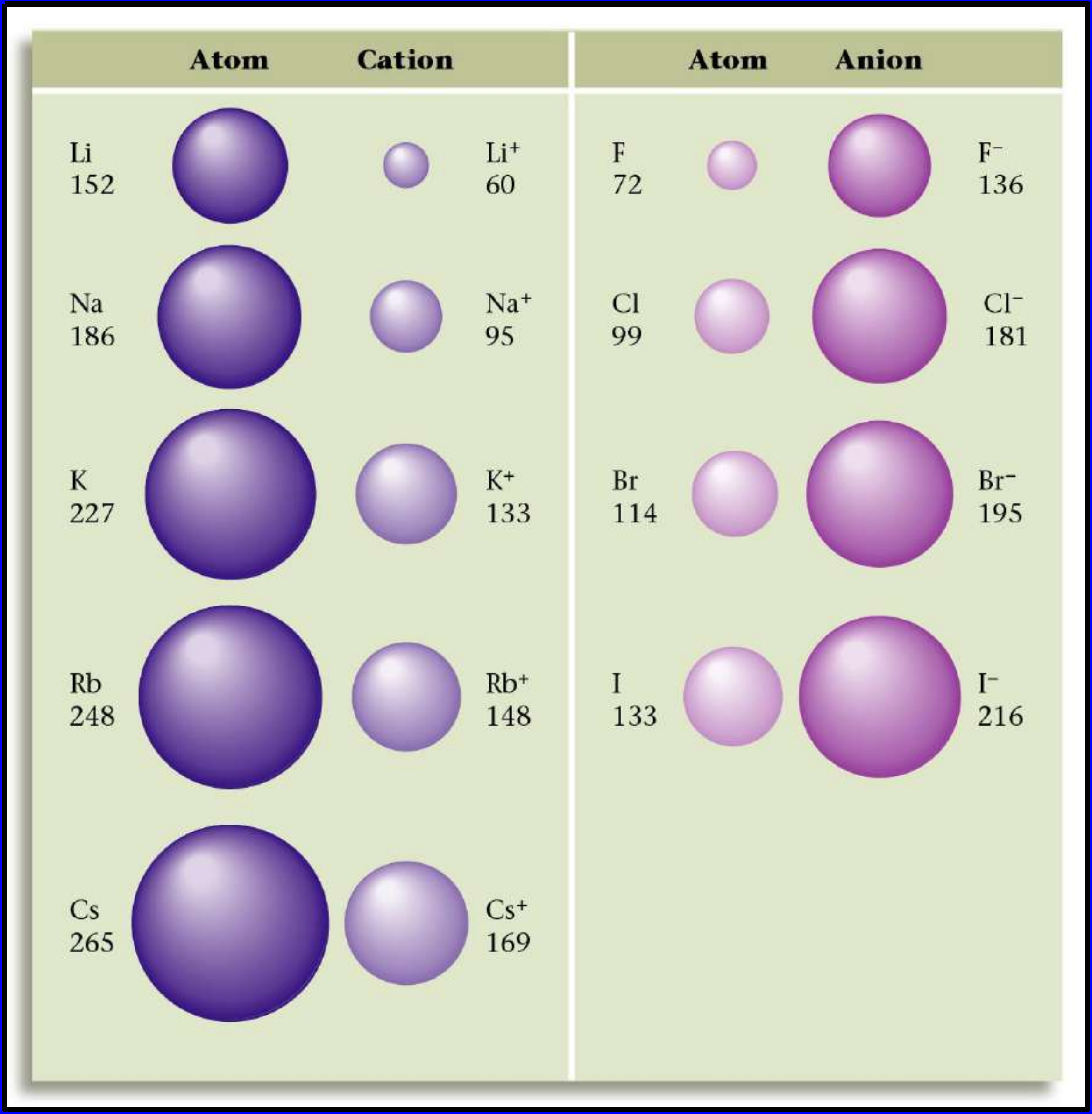


Figure 12.9: Relative sizes of some ions and their parent atoms.

Cation – always smaller than parent atom

Anion – always larger than parent atom



Polyatomic Ions

- Atoms in ion are held together by covalent bonds
- All atoms behave as one unit

Lewis Structures

- Bonding involves just valence electrons
- ***Lewis Structure***: representation of a molecule that shows how the valence electrons are arranged among the atoms in the molecule
- Named after G.N. Lewis – came up with idea while lecturing Chemistry class in 1902

Lewis Structures

- Only include valence electrons
- Use dots to represent electrons
- Hydrogen and Helium follow *duet rule*
- *Octet Rule:* eight electrons required – atoms can share
  - *Bonding pair* = shared electrons
  - *Lone pairs/unshared pairs*: electrons not involved in bonding

Steps for Writing Lewis Structures

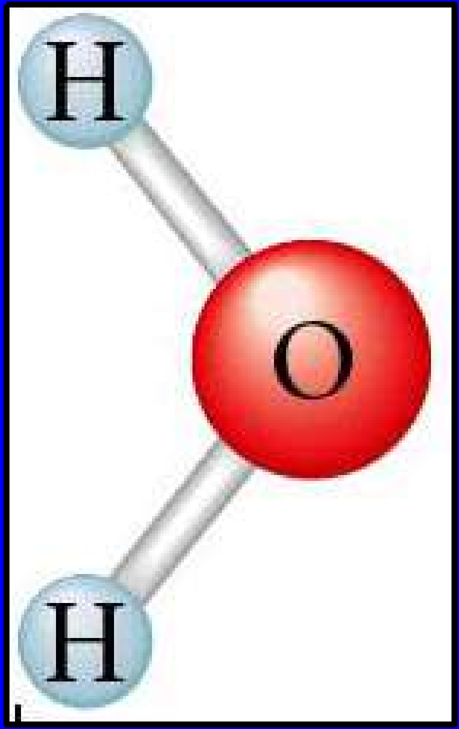
- Obtain sum of valence electrons from all atoms.
- Use one pair of electrons to form bond between each pair of atoms (can use line to represent 2 bonding electrons instead of dots)
- Arrange remaining electrons to satisfy duet or octet rule (there are exceptions to the octet rule)

Multiple Bonds

- *Single bond*: 2 atoms sharing one electron pair
- *Double bond*: 2 atoms sharing two pairs of electrons
- *Triple bond*: three electron pairs are shared
- *Resonance*: more than one Lewis structure can be drawn for the molecule
- Insert multiple bonds to satisfy octet rule

Molecular Structure

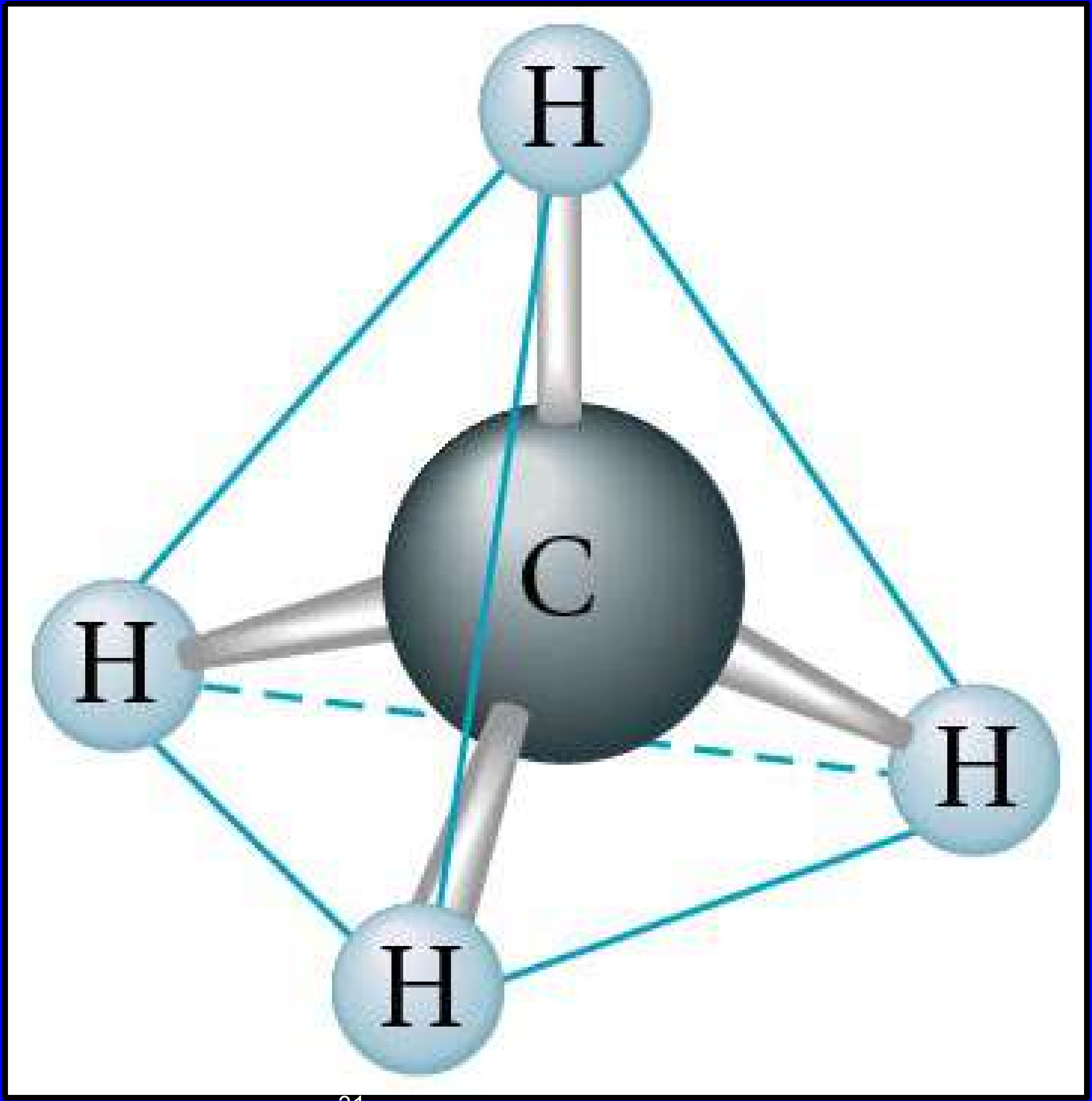
- aka Geometric Structure
- 3-D arrangement of the atoms in a molecule
- Example: water molecule
  - Bent or V-shaped
  - Describe precisely using bond angle = 105° for H_2O



Other Molecular Structures

- *Linear structure*: all atoms are in a line
  - Example: Carbon dioxide (see pg. 382)
  - Bond angle = 180°
- *Trigonal Planar*: triangular, planar with 120° bond angles
  - Example: BF_3 (see pg. 382)
- *Tetrahedral structure*: tetrahedron
  - Example: methane
  - Four identical triangular faces

Figure 12.12: Molecular structure of methane.



Molecular Structure: The VSEPR Model

- Structure very important to molecular properties
  - Determines taste
  - Biological molecules – structure change can convert cell from normal to cancerous
- Experimental methods exist for determining 3-D structure
- Useful to predict approximate structure

VSEPR Model

- Valence Shell Electron Pair Repulsion (VSEPR) model
- Used to predict molecular structure of molecules formed from nonmetals
- The structure around a given atom is determined by minimizing repulsions between electron pairs
- Bonding & non-bonding electron pairs around an atom are positioned as far apart as possible

VSEPR Model Rules (see ex pp 384-386)

- Two pairs of electrons on a central atom of a molecule are always placed at an angle of 180° to each other to give a linear arrangement
- Three pairs of electrons on a central atom in a molecule are always placed 120° apart in the same plane as the central atom – trigonal planar
- Four pairs of electrons on a central atom are always placed 109.5° apart - tetrahedral
- When every pair of electrons on central atom is shared, the molecular structure has same name as the arrangement of electron pairs
  - 2 = linear, 3 = trigonal planar, 4 = tetrahedral
- When one or more of the electron pairs around the central atom are unshared, the name of the structure is different from that for the arrangement of electron pairs (see table 12.4 # 4 & 5)

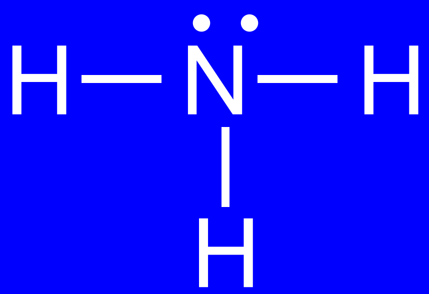
Steps for Predicting Molecular Structure Using VSEPR Model

- Draw the Lewis structure for the molecule
- Count electron pairs & arrange them to minimize repulsion (far apart)
- Determine position of atoms from the way the electron pairs are shared
- Determine name of molecular structure from the positions of atoms

Example 12.5: Predicting Molecular Structure using VSEPR

Predict structure of ammonia, NH_3

- Draw Lewis Structure

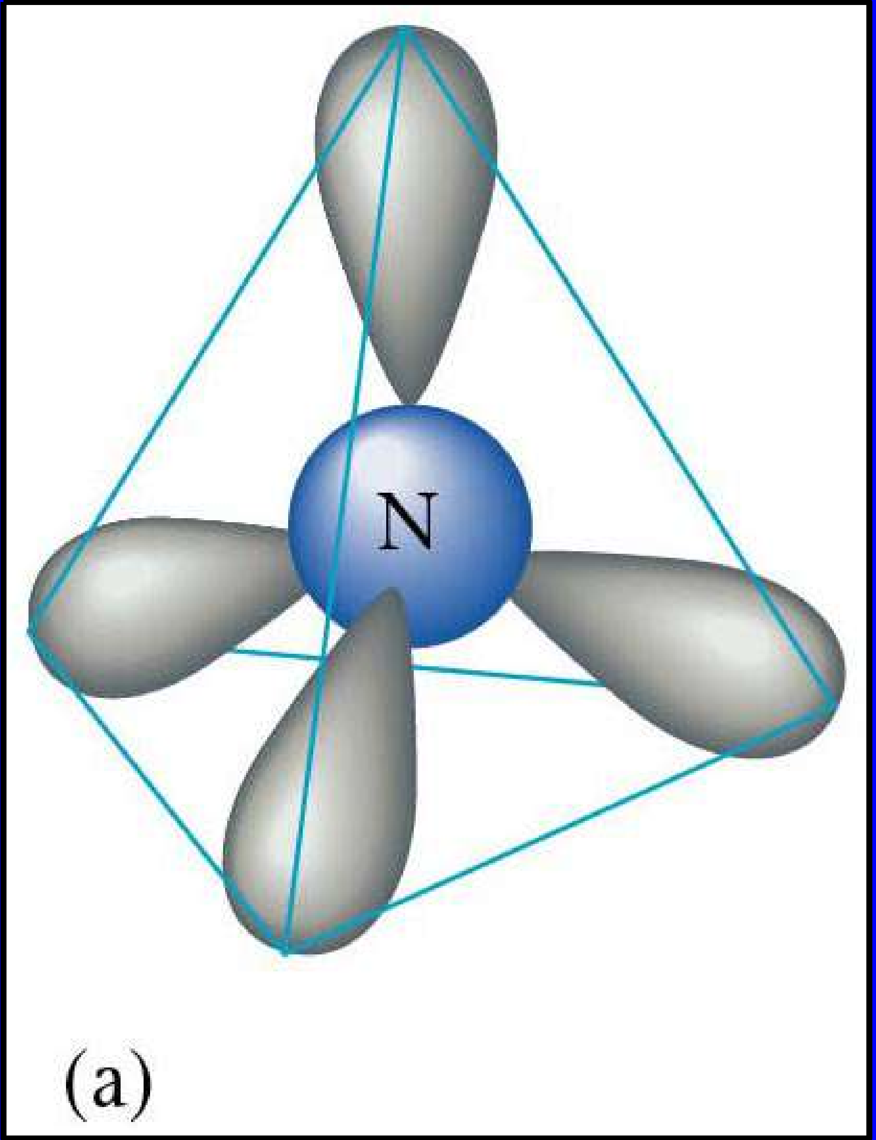


- Count pairs of electrons & arrange them to minimize repulsions (see next)

Example 12.5:

NH_3 has four pairs of electrons around the N atom (3 bonding) – best arrangement for 4 pairs is tetrahedral

Figure 12.13: Tetrahedral arrangement of electron pairs.



Step 3: Determine the positions of the atoms

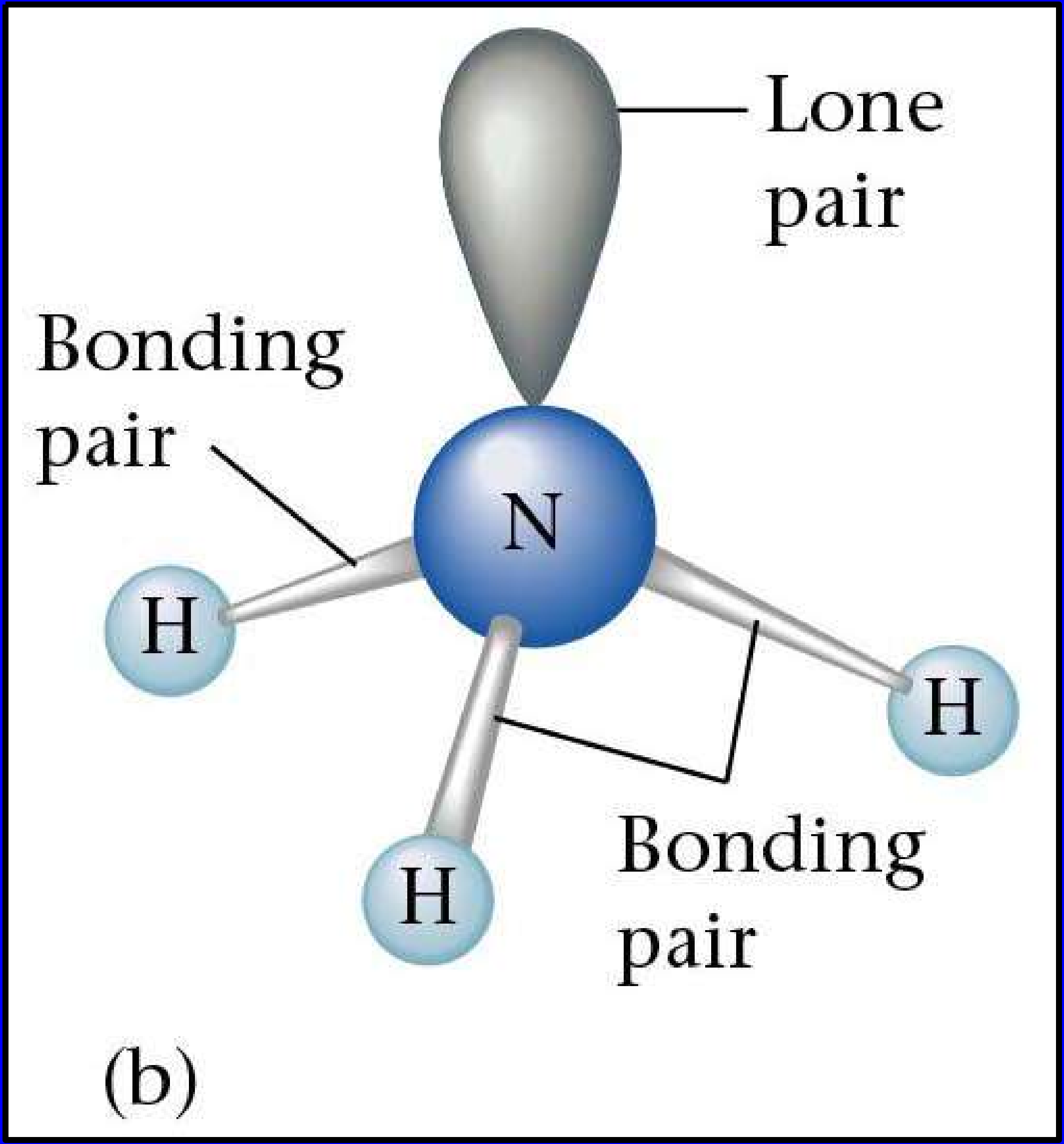
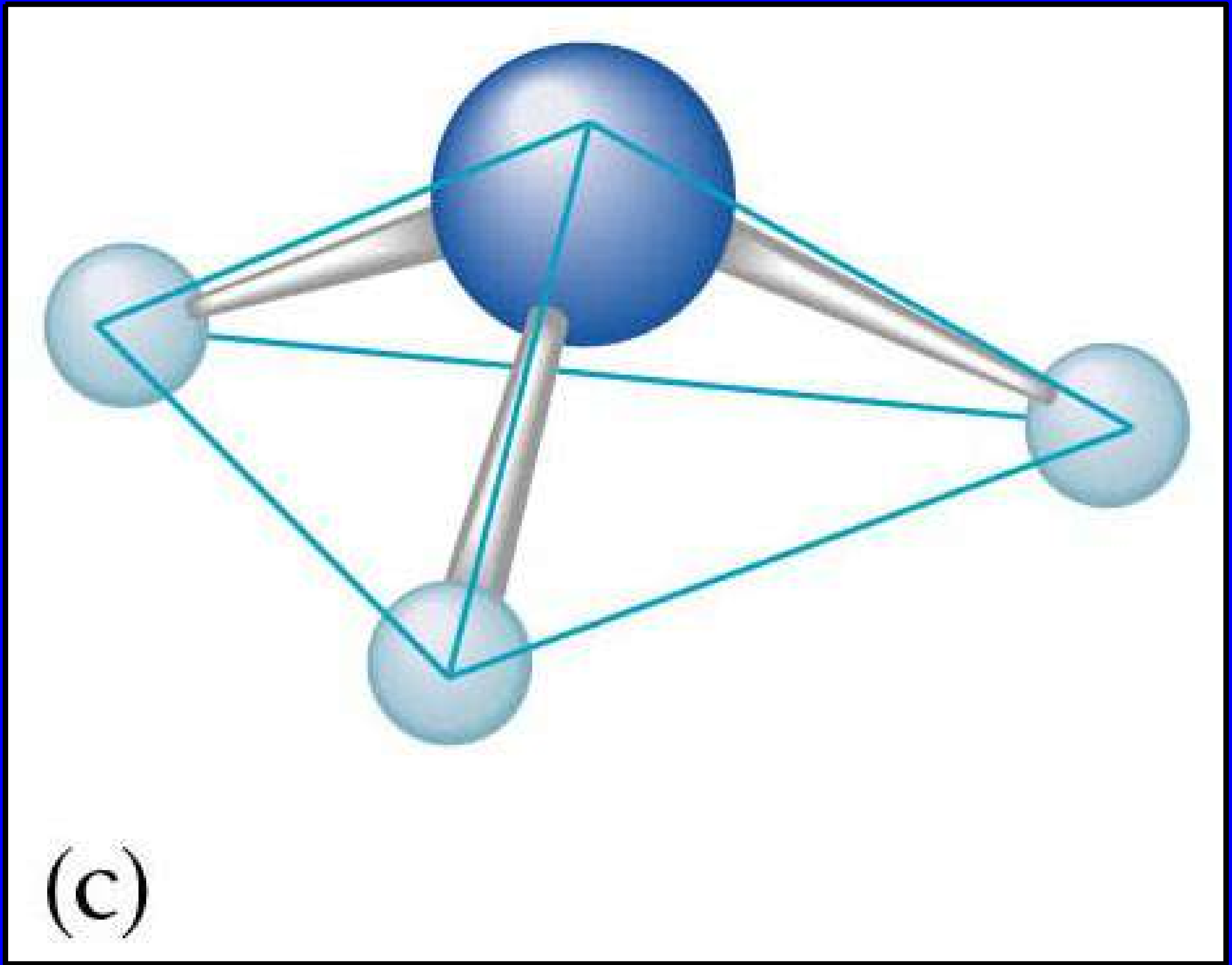


Figure 12.13: Hydrogen atoms occupy only three corners of the tetrahedron.

Step 4: Determine name of structure

- Name based on positions of the atoms
- Placement of electron pairs determines the structure, but name based on positions of atoms
- NH_3 has tetrahedral arrangement of electron pairs, but is not tetrahedral
- Structure is trigonal pyramid (one side different from other three)

Figure 12.13: The NH_3 molecule has the trigonal pyramid structure.



Example 12.6: describe molecular structure of water

Step 1: draw Lewis structure

Step 2: Count electron pairs & arrange to minimize repulsions

Figure 12.14: Tetrahedral arrangement of four electron pairs around oxygen.

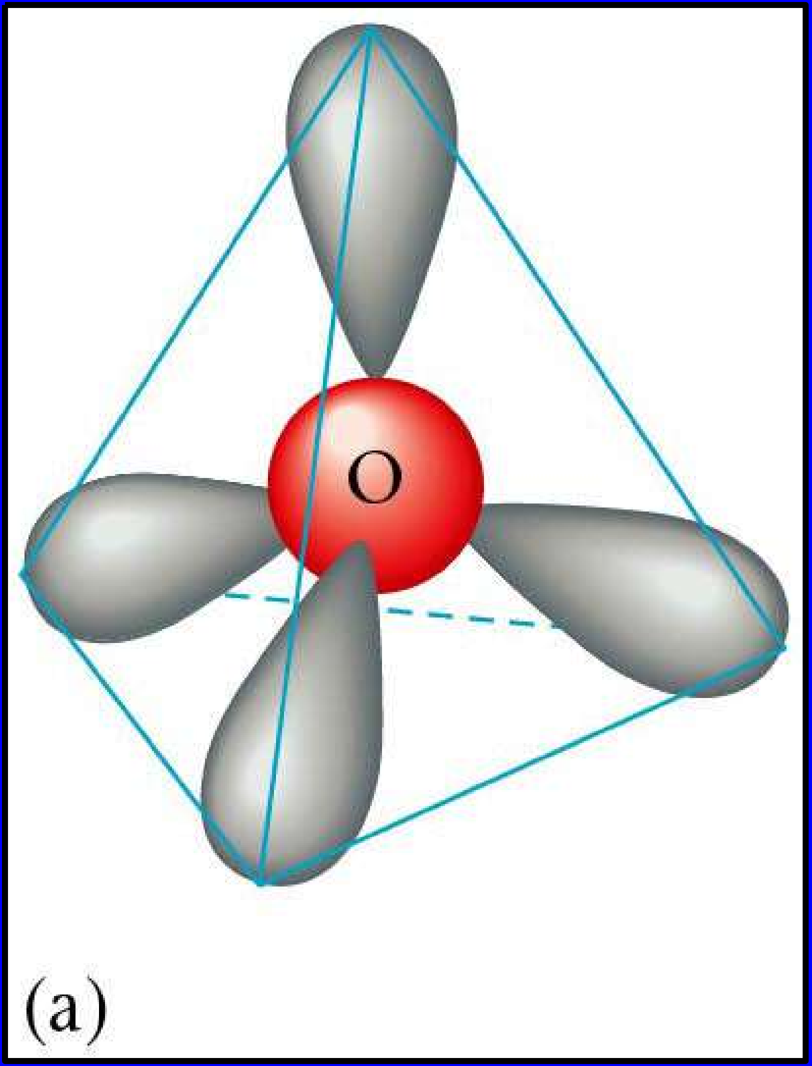


Figure 12.14: Two electron pairs shared between oxygen and hydrogen atoms.

Look at only atoms to determine structure

Step 3: tetrahedral arrangement of electron pairs, but not atoms – atoms form V-shape

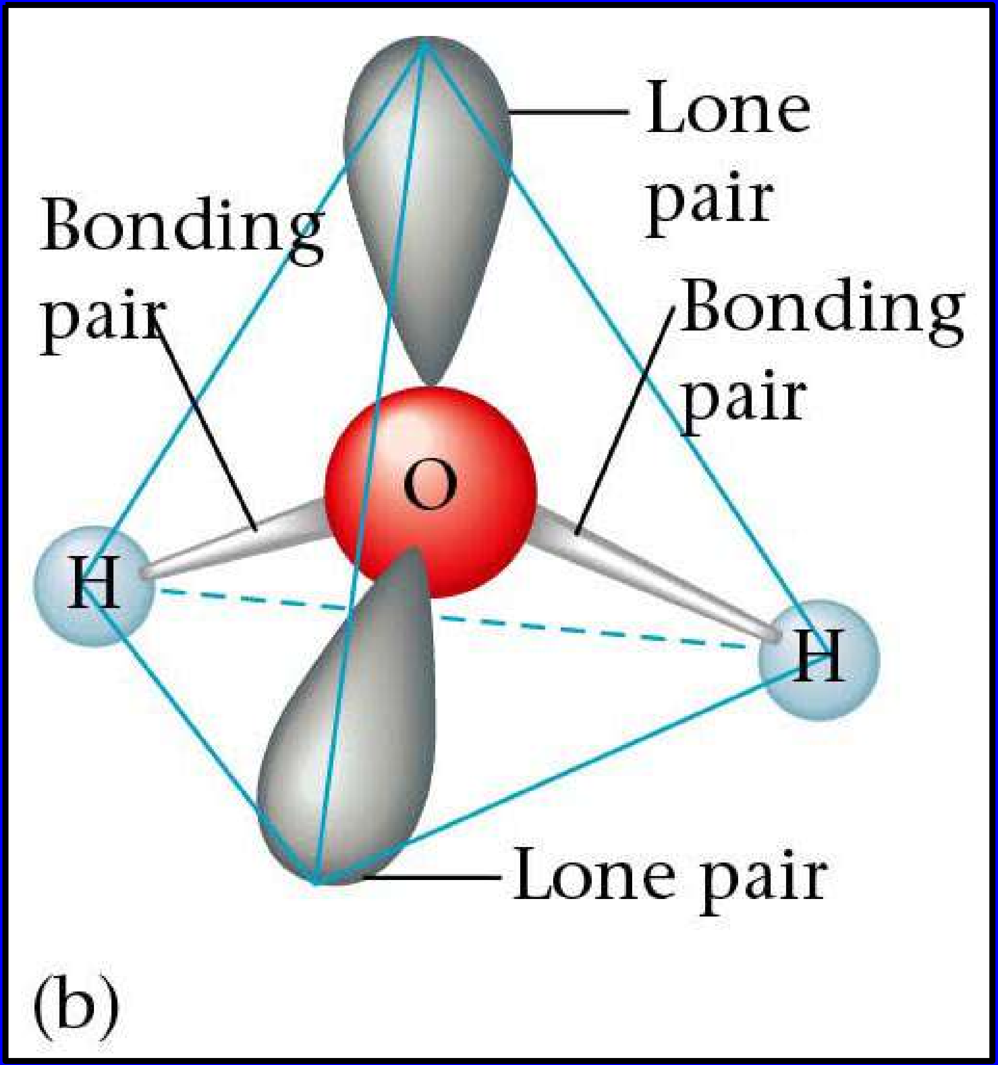


Figure 12.14: V-shaped molecular structure of the water molecule.

Step 4: Molecule is V-shaped or bent

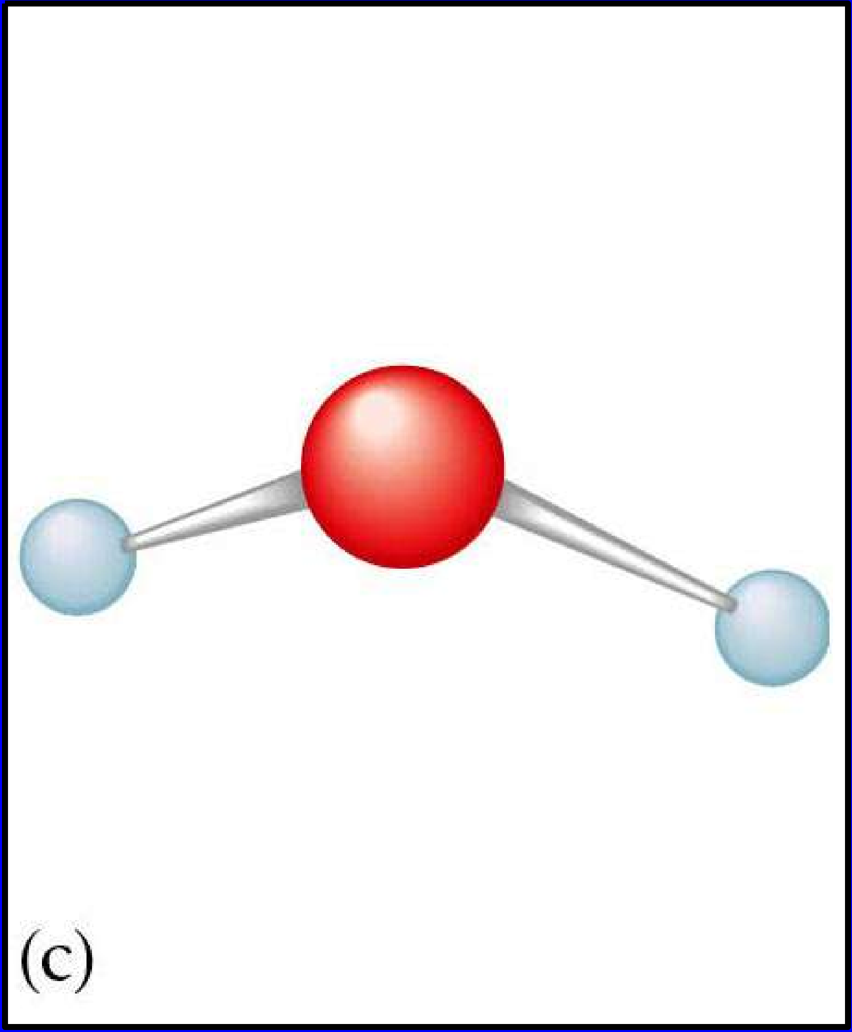
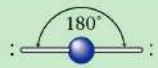
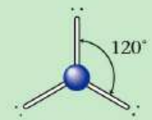
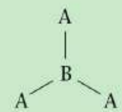
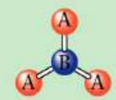
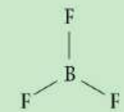
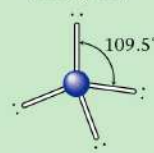
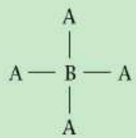
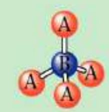
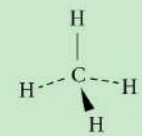
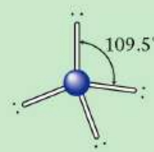
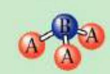
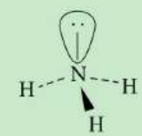
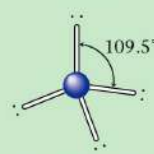
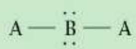
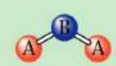
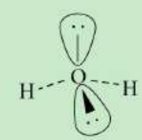


Table 12.4

TABLE 12.4

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

Case	Number of Electron Pairs	Bonds	Electron Pair Arrangement	Partial Lewis Structure	Molecular Structure	Example
1	2	2	 Linear	A — B — A	 Linear	F — Be — F BeF_2
2	3	3	 Trigonal planar (triangular)		 Trigonal planar (triangular)	 BF_3
3	4	4	 Tetrahedral		 Tetrahedral	 CH_4
4	4	3	 Tetrahedral		 Trigonal pyramid	 NH_3
5	4	2	 Tetrahedral		 Bent or V-shaped	 H_2O

Molecular Structure Involving Double Bonds

- When using the VSEPR model to predict the molecular geometry of a molecule, a double bond is counted the same as a single electron pair
- Four electrons involved in double bond do not act as two independent pairs, but are “tied together” for form one effective repulsive unit