

Chemical Bonding

Chapter 11

Structure Determines Properties!

- A cardinal principle of chemistry is that the macroscopic observed properties of a material are related to its microscopic structure.
 - And visa versa
- The microscopic structure entails
 - the kinds of atoms
 - the manner in which they are attached
 - their relationship to other molecules, like and dislike
 - the shape of the molecule

Chemical Bonds

- Forces that hold atoms together
- **Ionic bonds** are the forces of attraction between ions
 - ions formed by electron transfer
 - electrostatic forces
- **Covalent bonds** are the forces of attraction between two atoms which are sharing electrons

Ionic Bonds

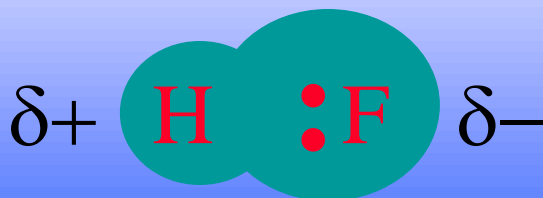
- Results from reaction between Metal and Nonmetal
- Metal loses electrons to form cation, Nonmetal gains electrons to form anion
- Ionic bond is the attraction between a positive ion and negative ion
- Larger Charge = Stronger Attraction
- Smaller Ion = Stronger Attraction
- ***No bond is 100% ionic!!***
- Electrostatic attraction nondirectional
 - no direct anion-cation pair, **No ionic molecule**
 - chemical formula is empirical formula, simply giving the ratio of ions based on charge balance
- Ions arranged in a pattern called a **crystal lattice**
 - maximizes attractions between + and - ions

Covalent Bonds

- Typical of molecular substances
- Atoms bond together to form molecules
 - strong attraction
- **Sharing pairs of electrons**
- Molecules attracted to each other weakly
- Often found between nonmetal atoms

Bond Polarity

- Covalent bonding between unlike atoms results in unequal sharing of the electrons
 - One end of the bond has larger electron density than the other
- The result is **bond polarity**
 - The end with the larger electron density gets a partial negative charge
 - The end that is electron deficient gets a partial positive charge



Electronegativity

- Measure of the ability of an atom to attract shared electrons
 - Larger electronegativity means atom attracts more strongly
 - Values 0.7 to 4.0
- Increases across period (left to right) on Periodic Table
- Decreases down group (top to bottom) on Periodic Table
- Larger difference in electronegativities means more polar bond
 - negative end toward more electronegative atom

Increasing electronegativity

H
2.1

Decreasing electronegativity

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

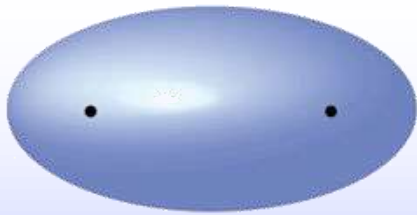
Dipole Moment

- Bond polarity results in an unequal electron distribution, resulting in areas of partial positive and partial negative charge
- Any molecule that has a center of positive charge and a center of negative charge in different points is said to have a **dipole moment**
- If a molecule has more than one polar covalent bond, the areas of partial negative and positive charge for each bond will partially add to or cancel out each other
- The end result will be a molecule with one center of positive charge and one center of negative charge
- The dipole moment affects the attractive forces between molecules and therefore the physical properties of the substance

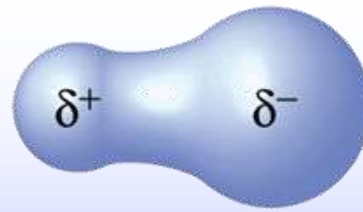
Table 11.1 The Relationship Between Electronegativity and Bond Type

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

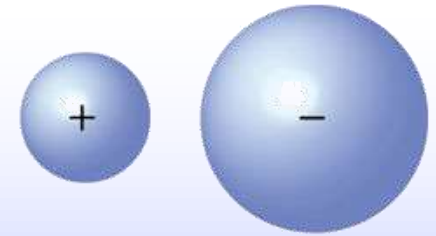
The three possible types of bonds: (a) a covalent bond formed between identical atoms; (b) a polar covalent bond, with both ionic and covalent components; and (c) an ionic bond, with no electron sharing.



(a)



(b)



(c)

Electron Arrangements And Ion Charge

- We know
 - Group 1A metals form ions with +1 charge
 - Group 2A metals form ions with +2 charge
 - Group 7A nonmetals form ions with -1 charge
 - Group 6A nonmetals form ions with -2 charge
 - Group 8A nonmetals do not form ions, in fact they are extremely unreactive

Electron Arrangements and Ion Charge

- Representative Metals form cations by losing enough electrons to get the same electron configuration as the previous noble gas
- Nonmetals form anions by gaining enough electrons to get the same electron configuration as the next noble gas

Atom	Atoms Electron Config	Ion	Ions Electron Config
Na	[Ne]3s ¹	Na ⁺¹	[Ne]
Mg	[Ne]3s ²	Mg ⁺²	[Ne]
Al	[Ne]3s ² 3p ¹	Al ⁺³	[Ne]
O	[He]2s ² 2p ⁴	O ⁻²	[Ne]
F	[He]2s ² 2p ⁵	F ⁻¹	[Ne]

Electron Arrangements and Ionic Bonding

- Representative metals lose their valence electrons to form cations
- Nonmetals gain electrons so their valence shell has the same electron arrangement as the next noble gas
- There have to be enough electrons from the metals atoms to supply the needed electrons for the nonmetal atoms
 - Allows us to predict the formulas of ionic compounds
- In Polyatomic ions, the atoms in the ion are connected with covalent bonds. The ions are attracted to oppositely charged ions to form an ionic compound

Properties of Ionic Compounds

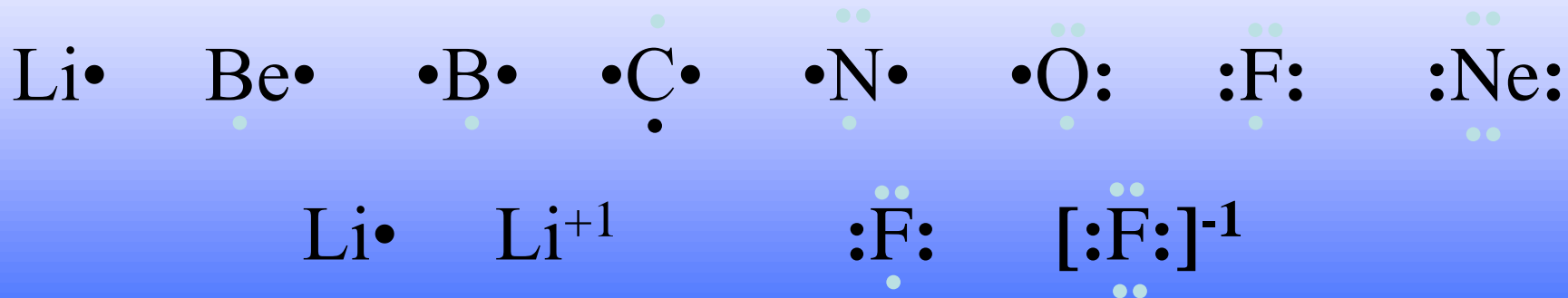
- All solids at room temperature
 - Melting points greater than 300°C
- Liquid state conducts electricity, solid state does not
 - Liquid = molten
- Brittle and Hard
- Often soluble in water, and when dissolved the solution becomes an electrical conductor
 - When ionic compounds containing polyatomic ions dissolve, the covalent bonds holding the polyatomic ion do not break, the ion stays together even though it separates from the other ion
 - All strong electrolytes

Bonding and Structure of Ionic Compounds

- **Crystal Lattice** = geometric pattern determined by the size and charge of the ions
- Anions larger than cation
 - Almost always
 - Anions larger than parent atom, Cations smaller than parent atom
- Anions generally considered “hard” spheres packed as efficiently as possible, with the cations occupying the “holes” in the packing
- Arrangement results in each cation being surrounded by as many anions as will fit
 - And visa versa
 - **Maximizes attractions between ions**

Lewis Symbols of Atoms and Ions

- Also known as electron dot symbols
- Use symbol of element to represent nucleus and inner electrons
- Use dots around the symbol to represent valence electrons
 - put one electron on each side first, then pair
- Elements in the same group have the same Lewis symbol
 - Because they have the same number of valence electrons
- Cations have Lewis symbols without valence electrons
- Anions have Lewis symbols with 8 valence electrons



Writing Lewis Structures of Molecules

- Count the *total* number of valence electrons from all the atoms
- Attach the atoms together with one pair of electrons
 - A line is often used as shorthand for a pair of electrons that attach atoms together
- Arrange the remaining electrons in pairs so that all hydrogen atoms have 2 electrons (1 bond) and other atoms have 8 electrons (combination of bonding and nonbonding)
- Occasionally atoms may violate this rule
 - Nonbonding pairs of electrons are also known as Lone Pairs

Covalent Bonds

- Single Covalent Bond the atoms share 2 electrons,
 - (1 pair)
- Double Covalent Bond the atoms share 4 electrons,
 - (2 pairs)
- Triple Covalent Bond the atoms share 6 electrons,
 - (3 pairs)
- Bond Strength = Triple > Double > Single
 - For bonds between same atoms, $C\equiv N > C=N > C-N$
 - Though Double not 2x the strength of Single and Triple not 3x the strength of Single
- Bond Length = Single > Double > Triple
 - For bonds between same atoms, $C-N > C=N > C\equiv N$

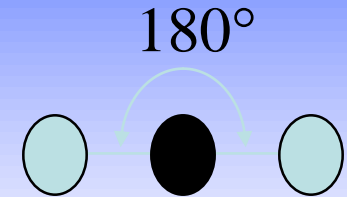
Problems with Lewis Structures

- Some atoms do not tend to follow the octet rule
 - B and Be often found octet deficient
 - Elements in the 3rd Period or below often have expanded octets
- Some molecules have an odd number of electrons
- Impossible to accurately draw Lewis structure of molecules that exhibit resonance
- Sometimes the Lewis Structure does not accurately describe a structure that explains all the observed properties of the molecule
- The paramagnetic behavior of O₂

Some Geometric Figures

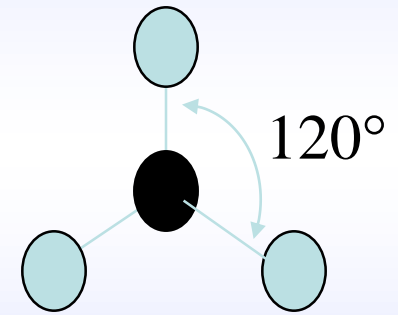
- **Linear**

- 2 atoms on opposite sides of central atom
- 180° bond angles



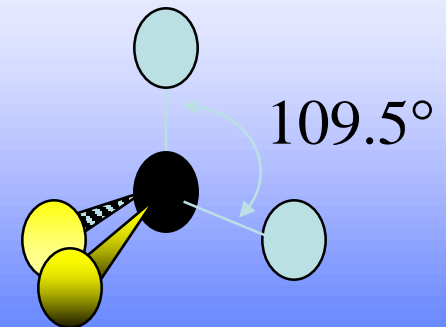
- **Trigonal Planar**

- 3 atoms form a triangle around the central atom
- Planar
- 120° bond angles



- **Tetrahedral**

- 4 surrounding atoms form a tetrahedron around the central atom
- 109.5° bond angles



Predicting Molecular Geometry

- VSEPR Theory
 - Valence Shell Electron Pair Repulsion
- The shape around the central atom(s) can be predicted by assuming that the areas of electrons on the central atom will repel each other
- Each Bond counts as 1 area of electrons
 - single, double or triple all count as 1 area
- Each Lone Pair counts a 1 area of electrons
 - Even though lone pairs are not attached to other atoms, they do “occupy space” around the central atom
 - Lone pairs generally “push harder” than bonding electrons, affecting the bond angle

Shapes

- Linear
 - 2 areas of electrons around the central atom, both bonding
 - Or two atom molecule as trivial case
- Trigonal
 - 3 areas of electrons around the central atom
 - All Bonding = **trigonal planar**
 - 2 Bonding + 1 Lone Pair = **trigonal bent**
- Tetrahedral
 - 4 areas of electrons around the central atom
 - All Bonding = **tetrahedral**
 - 3 Bonding + 1 Lone Pair = **trigonal pyramid**
 - 2 Bonding + 2 Lone Pair = **tetrahedral bent or V-shaped**

The tetrahedral molecular structure of methane.

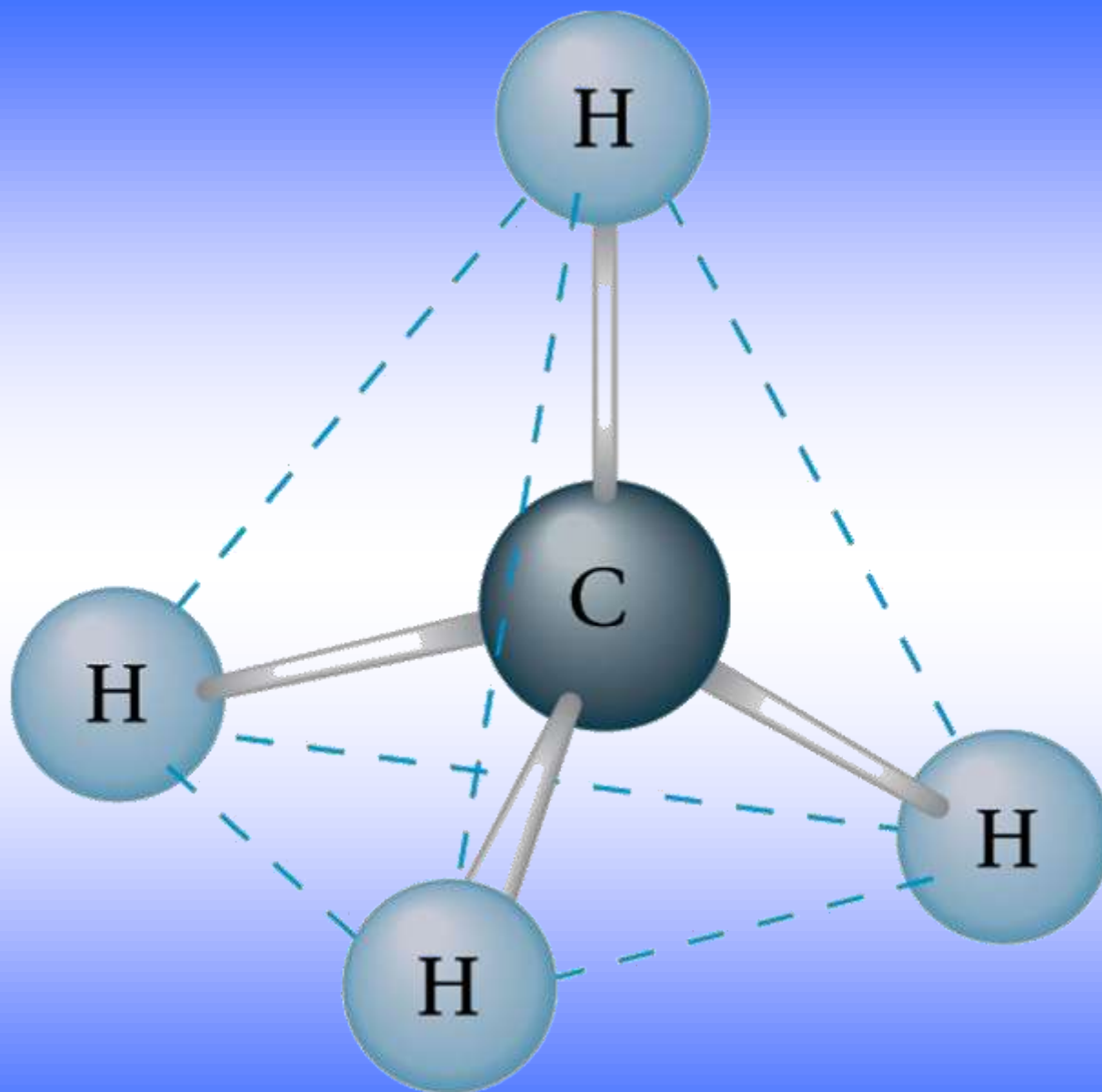
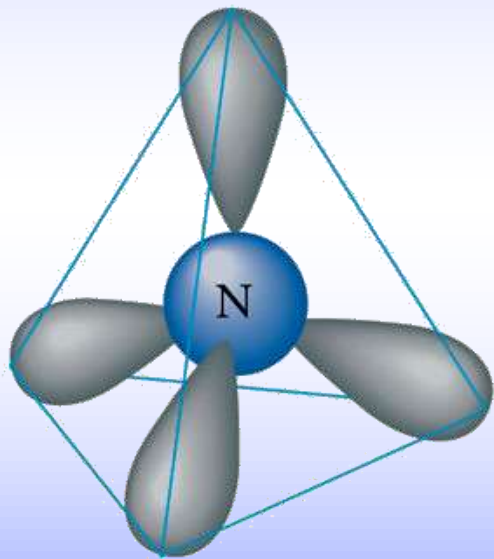


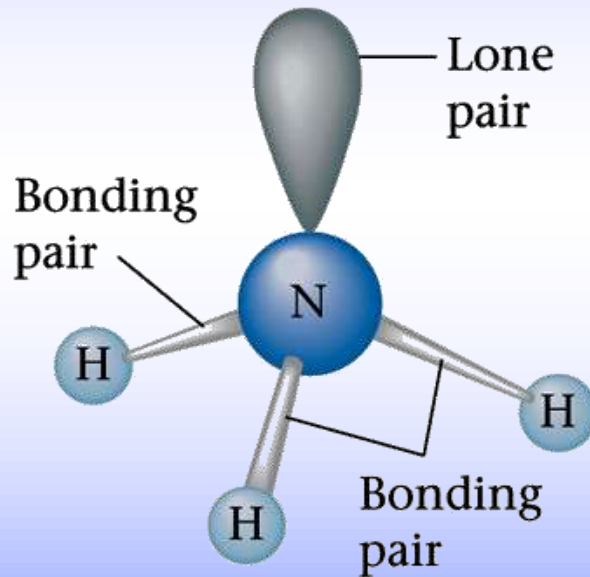
Table 11.4 Arrangements of Electron Pairs and the Resulting Molecular Structures for Two, Three, and Four Electron Pairs

Case	Number of Electron Pairs	Number of Bonds	Electron Pair Arrangement	Ball-and-Stick Model	Angle Between Pairs	Molecular Structure	Partial Lewis Structure	Ball-and-Stick Model	Example
1	2	2	Linear		180°	Linear	A—B—A		BeF ₂
2	3	3	Trigonal planar (triangular)		120°	Trigonal planar (triangular)			
3	4	4	Tetrahedral		109.5°	Tetrahedral			CH ₄
4	4	3	Tetrahedral		109.5°	Trigonal pyramid			NH ₃
5	4	2	Tetrahedral		109.5°	Bent or V-shaped			H ₂ O

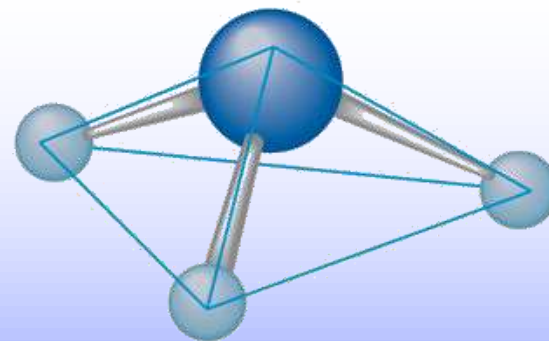
(a) the tetrahedral arrangement of electron pairs around the nitrogen atom in the ammonia molecule. (b) Three of the electron pairs around nitrogen are shared with the hydrogen atoms as shown, and one is a lone pair. (c) The NH_3 molecule has the trigonal pyramid structure.



(a)

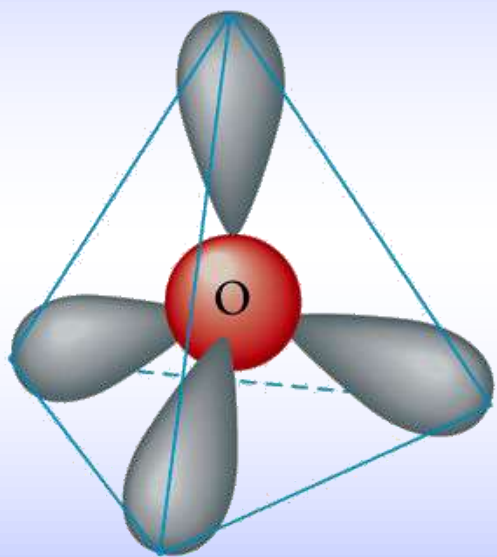


(b)

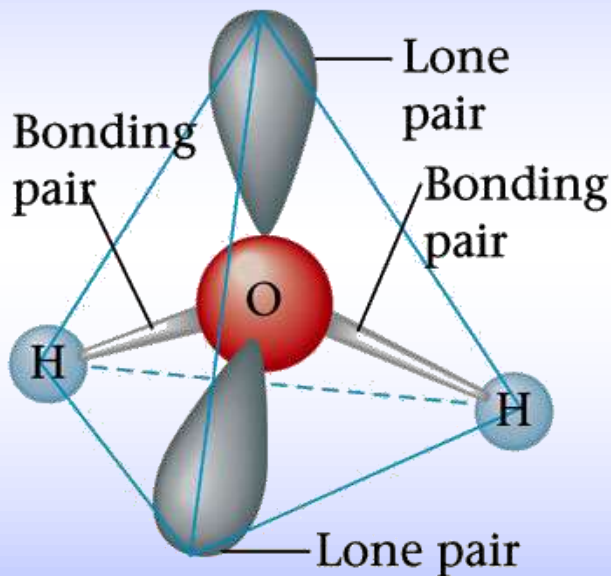


(c)

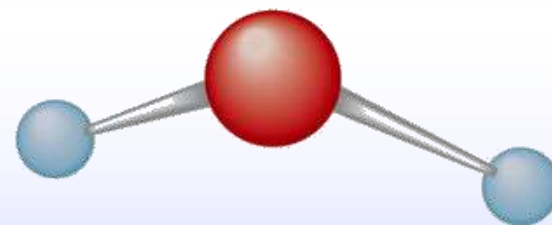
(a) The tetrahedral arrangement of the four electron pairs around oxygen in the water molecule. (b) Two of the electron pairs are shared between oxygen and the hydrogen atoms. (c) The V-shaped molecular structure of the water molecule.



(a)



(b)



(c)

When liquid oxygen is poured between the poles of a magnet, it “sticks” until it boils away.

