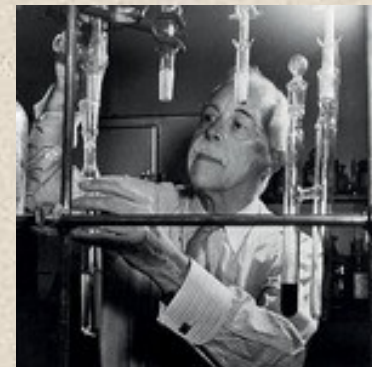
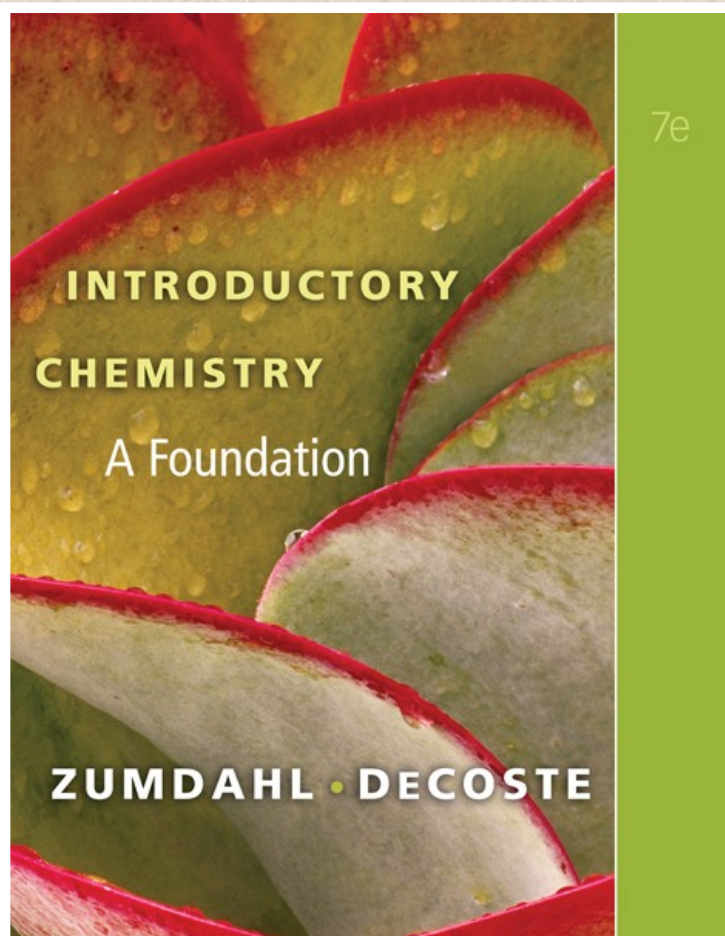




CHAPTER SIX

CHEMICAL BONDING

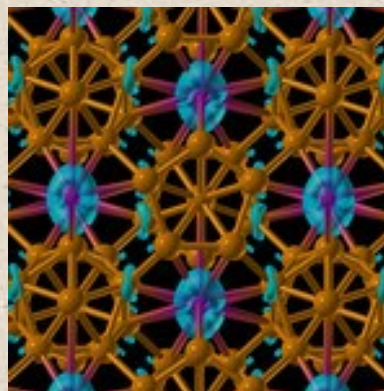


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- 6.4 Stable Electron Configurations and Charges on Ions
- 6.5 Ionic Bonding and Structures of Ionic Compounds
- 6.6 Lewis Structures
- 6.7 Lewis Structures of Molecules with Multiple Bonds
- 6.8 Molecular Structure
- 6.9 Molecular Structure: The VSEPR Model

Chapter 6

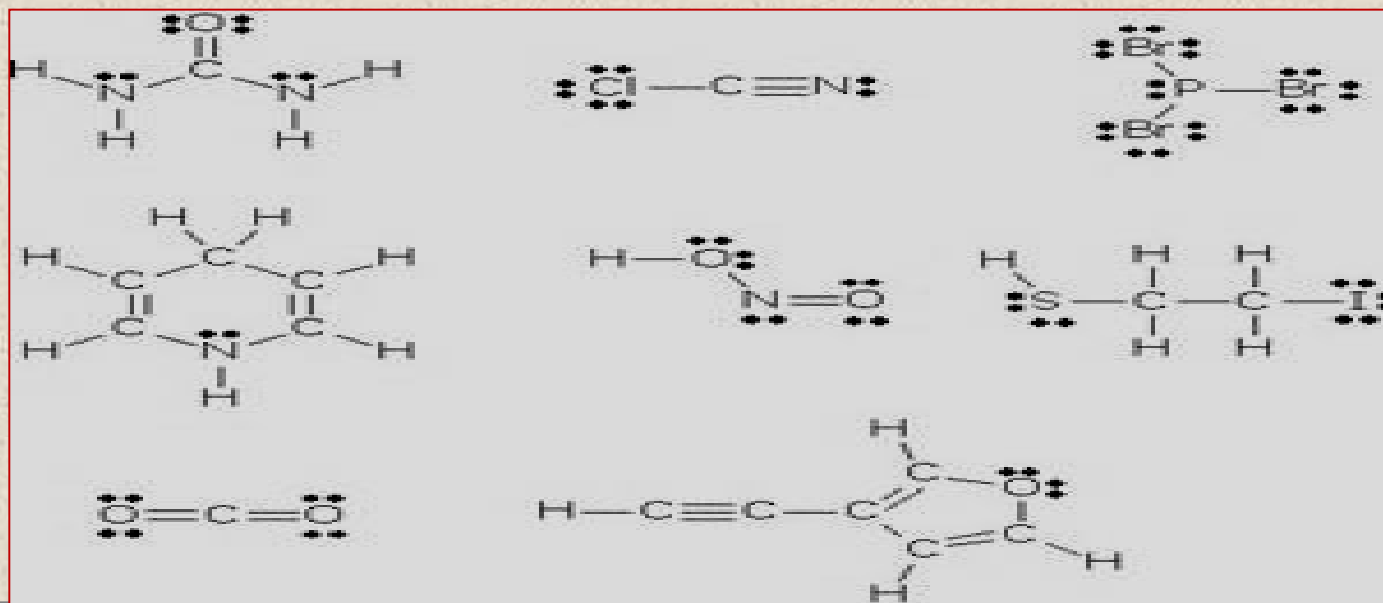
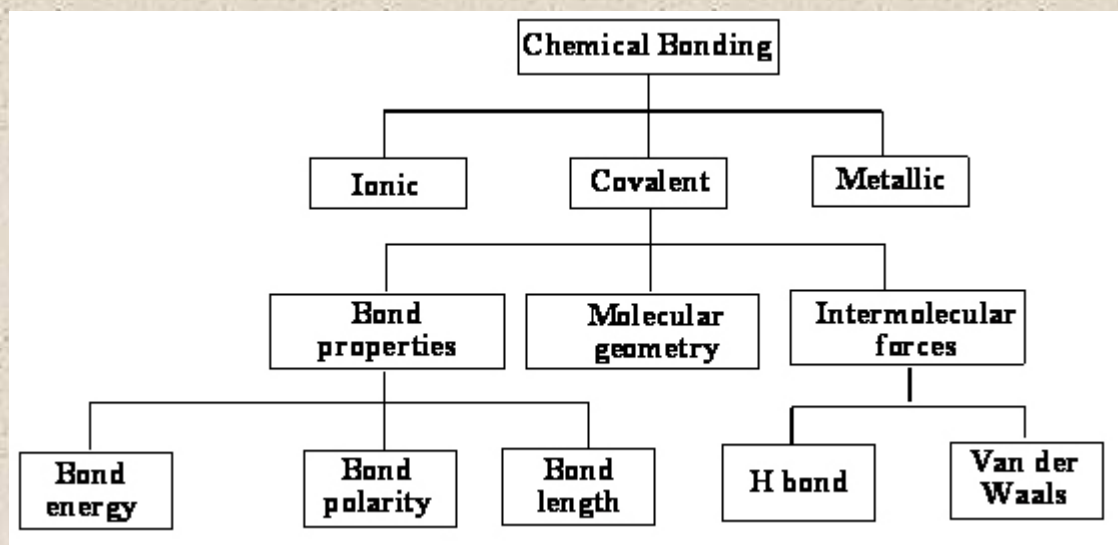


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Chemical Bond

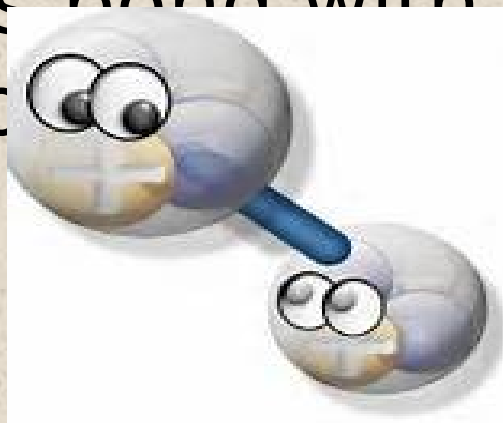
+ Bonding, the way atoms are attracted to each other to form molecules, determines nearly all of the chemical properties we see.

+ And, as we shall see, the number **“8”** is very important to chemical bonding.

8

Questions to Consider

- **What** is meant by the term “**chemical bond**”?
- **Why** do atoms bond with each other to form Compounds?
- **How** do atoms bond with each other to form compounds?



Section 6.1

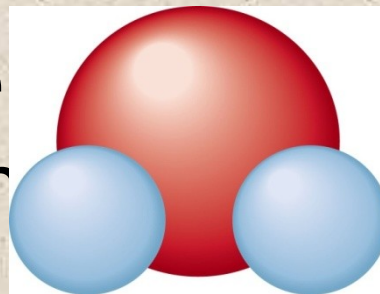
Types of Chemical Bonds

A Chemical Bond

 The forces that hold groups of atoms together and make them function as a unit.

 A bond will form if the energy of the aggregate is lower than that of the separated atoms.

 **Bond energy** – energy required to break a chemical bond

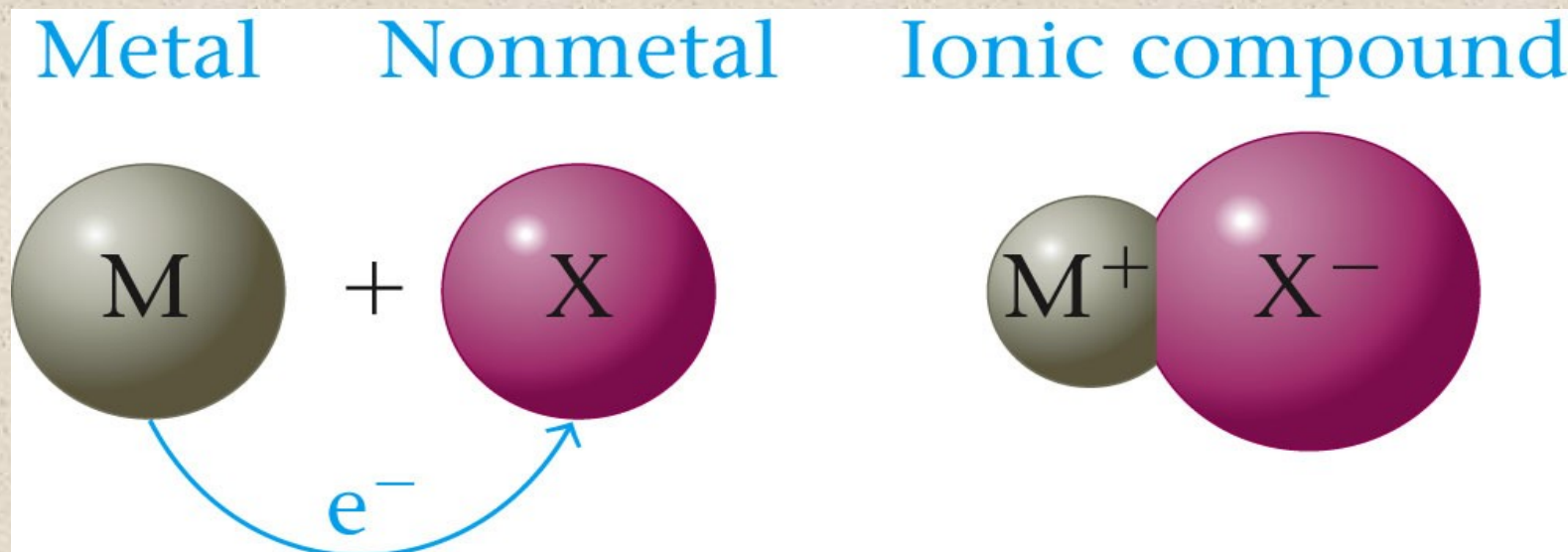


Section 6.1

Types of Chemical Bonds

Ionic Bonding

- Ionic compound results when a metal reacts with a nonmetal.
- Electrons are transferred.

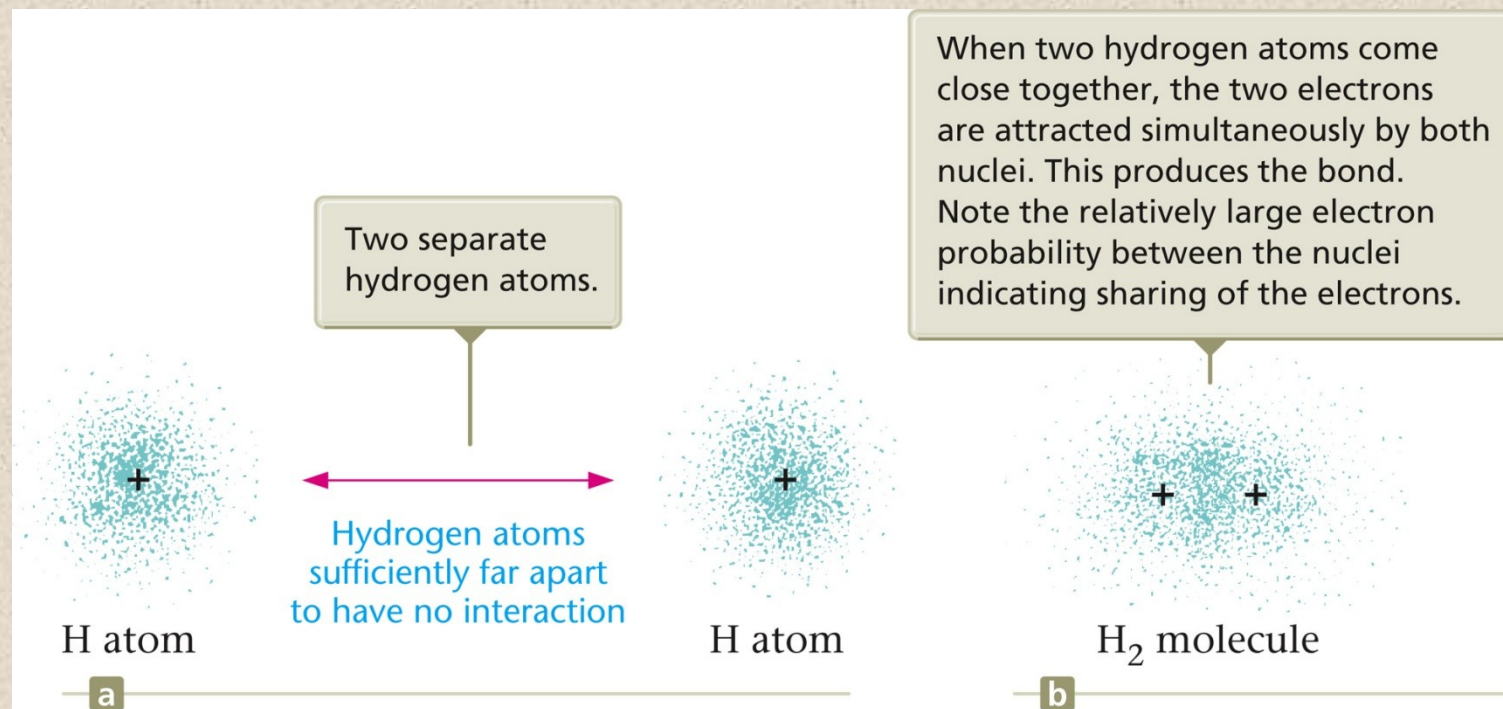


Section 6.1

Types of Chemical Bonds

Covalent Bonding

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



Section 6.1

Types of Chemical Bonds

Polar Covalent Bond

- Unequal sharing of electrons between atoms in a molecule.
- One atom attracts the electrons more than the other atom.

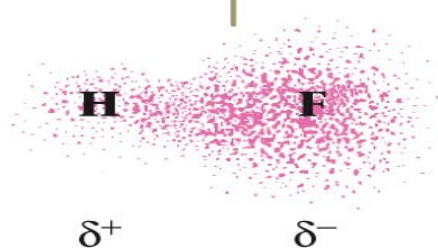
• Re
(p

What the probability map would like if the two electrons in the H—F bond were shared equally.



a

The actual situation, where the shared pair spends more time close to the fluorine atom than to the hydrogen atom. This gives the fluorine a slight excess of negative charge and the hydrogen a slight deficit of negative charge (a slight positive charge).

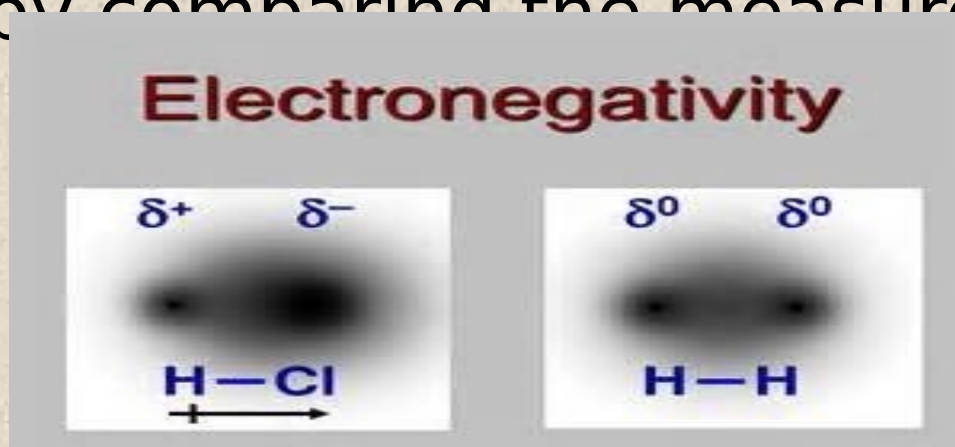


b

Section 6.2

Electronegativity

- ✚ The ability of an atom in a molecule to attract shared electrons to itself.
- ✚ For a molecule HX, the relative electronegativities of the H and X atoms are determined by comparing the measured H-X bond energy with the bond energy.



Section 6.2

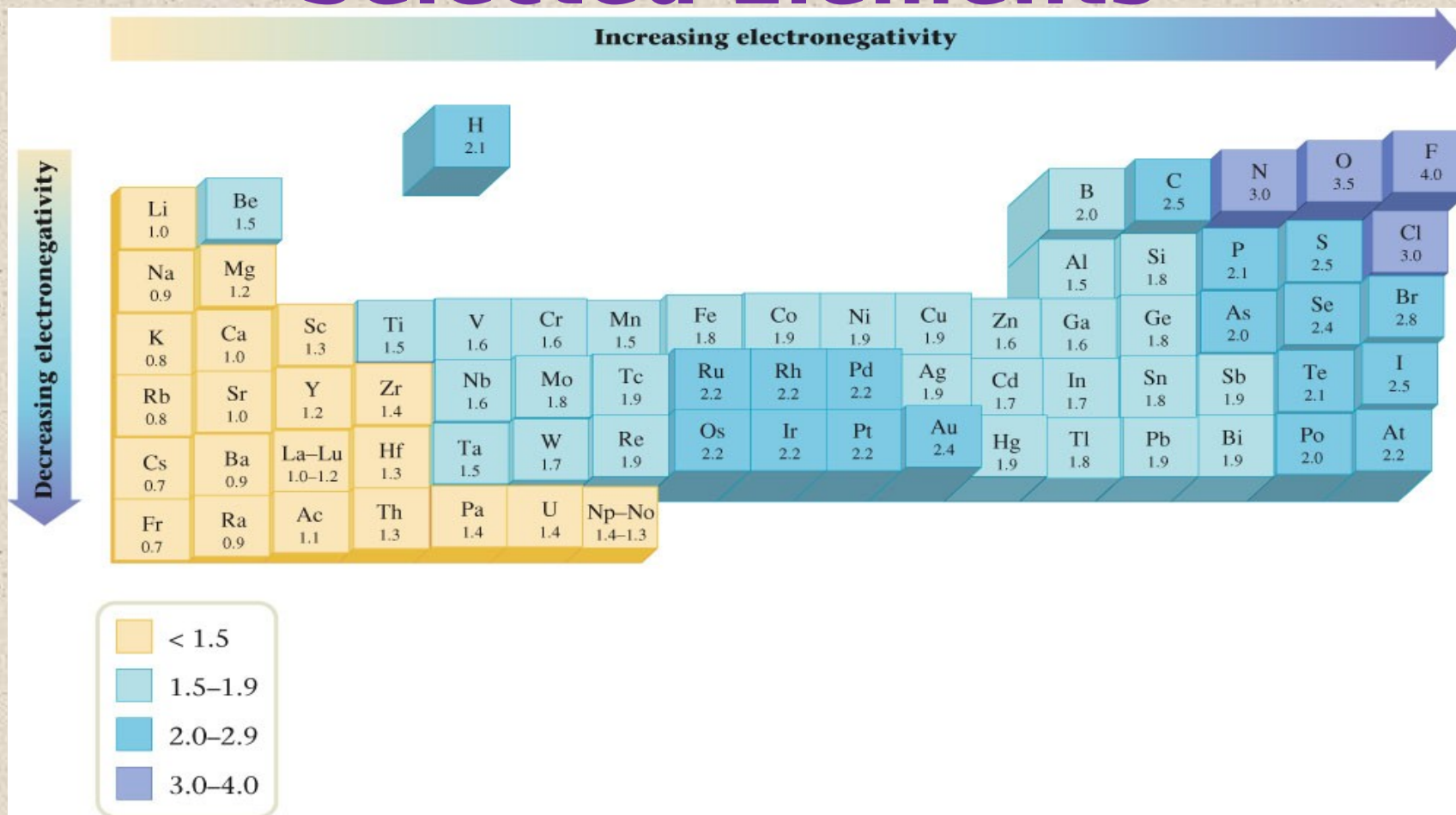
Electronegativity

- On the periodic table, electronegativity generally **increases** across a period and **decreases** down a group.
- The range of electronegativity values is from **4.0** for **fluorine** (the most electronegative) to 0.7 for cesium and francium (the least electronegative).

Section 6.2

Electronegativity

Electronegativity Values for Selected Elements



Section 6.2

Electronegativity



Concept Check

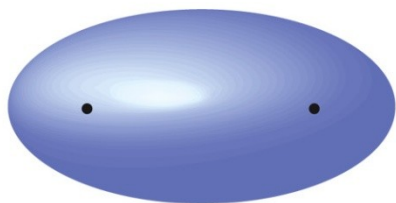
If lithium and fluorine react, which has **more attraction** for an electron? Why?

In a bond between fluorine and iodine, which has **more attraction** for an electron? Why?

Section 6.2

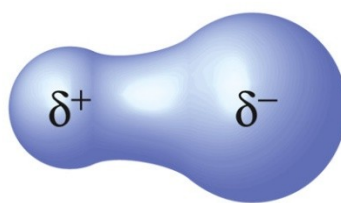
Electronegativity

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



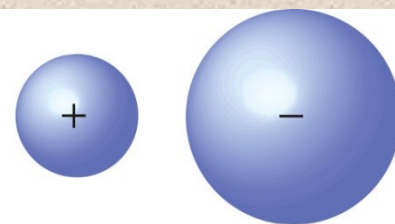
a

A covalent bond formed between identical atoms.



b

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



c

An ionic bond, with no electron sharing.

Section 6.2

Electronegativity

Using Electronegativity to Determine Bond Polarity

Using the electronegativity values given in Figure 12.4, arrange the following bonds in order of increasing polarity: H—H, O—H, Cl—H, S—H, and F—H.

Solution

The polarity of the bond increases as the difference in electronegativity increases. From the electronegativity values in Figure 12.4, the following variation in bond polarity is expected (the electronegativity value appears in parentheses below each element).

Bond	Electronegativity Value	Difference in Electronegativity Values	Bond Type	Polarity
H—H	(2.1) (2.1)	$2.1 - 2.1 = 0$	Covalent	
S—H	(2.5) (2.1)	$2.5 - 2.1 = 0.4$	Polar covalent	
Cl—H	(3.0) (2.1)	$3.0 - 2.1 = 0.9$	Polar covalent	
O—H	(3.5) (2.1)	$3.5 - 2.1 = 1.4$	Polar covalent	
F—H	(4.0) (2.1)	$4.0 - 2.1 = 1.9$	Polar covalent	

Therefore, in order of increasing polarity, we have

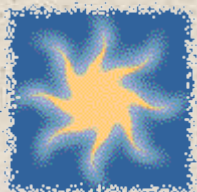
H—H S—H Cl—H O—H F—H

Least polar

Most polar

Section 6.2

Electronegativity



Exercise

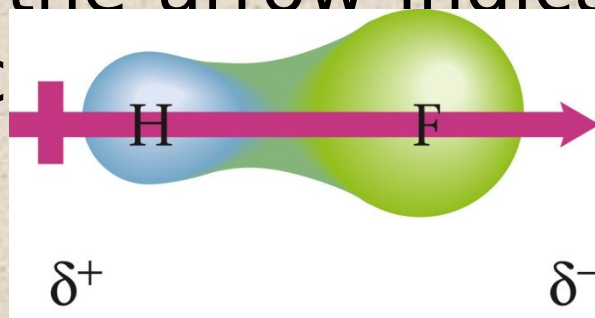
Arrange the following bonds from **most to least polar**:

- | | | | |
|----|-------|-------|-------|
| a) | N-F | O-F | C-F |
| b) | C-F | N-O | Si-F |
| c) | Cl-Cl | B-Cl | S-Cl |
| a) | C-F, | N-F, | O-F |
| b) | Si-F, | C-F, | N-O |
| c) | B-Cl, | S-Cl, | Cl-Cl |

Bond Polarity and Dipole Moments

Dipole Moment

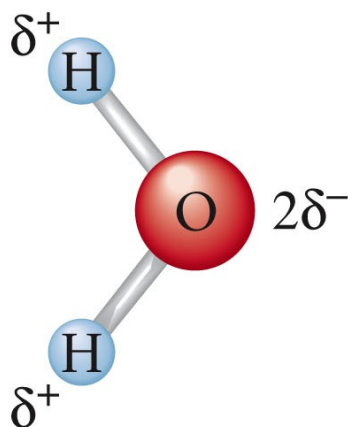
- Property of a molecule whose charge distribution can be represented by a center of positive charge and a center of negative charge.
- Use an arrow to represent a dipole moment.
 - Point to the negative charge center with the tail of the arrow indicating the positive center of charge.



Section 6.3

Bond Polarity and Dipole Moments

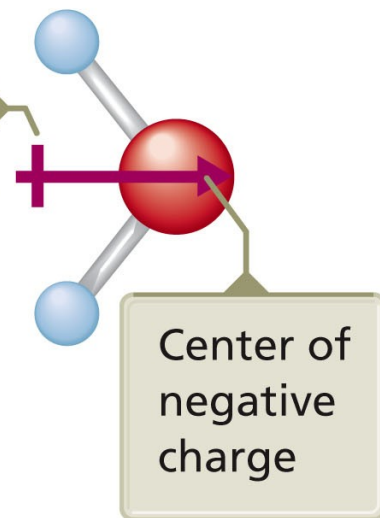
Dipole Moment in a Water Molecule



a

The charge distribution in the water molecule. The oxygen has a charge of $2\delta^-$ because it pulls δ^- of charge from each hydrogen atom ($\delta^- + \delta^- = 2\delta^-$).

Center of positive charge



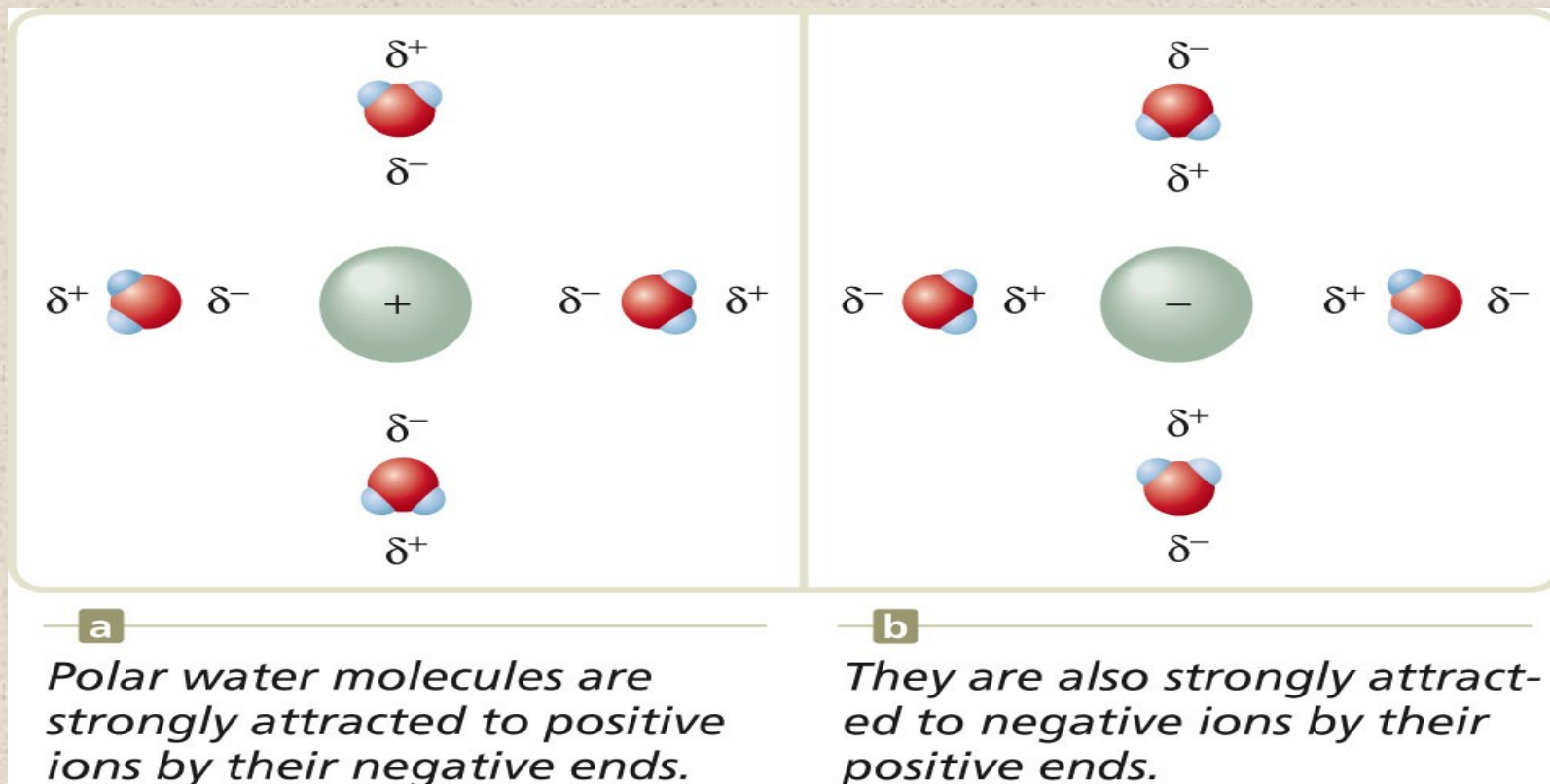
b

The water molecule behaves as if it had a positive end and a negative end, as indicated by the arrow.

Section 6.3

Bond Polarity and Dipole Moments

- The polarity of water affects its properties.
 - Permits ionic compounds to dissolve in it.



- Causes water to remain liquid at higher temperature.

6.4 Stable Electron Configurations and Charges on Ions

- Group 7 nonmetals form 1- anions.
- Group 6 elements always form 2- anions.

Section 6.4

Stable Electron Configurations and Charges on Ions

The Formation of Ions by Metals and Nonmetals

Table 12.2 The Formation of Ions by Metals and Nonmetals

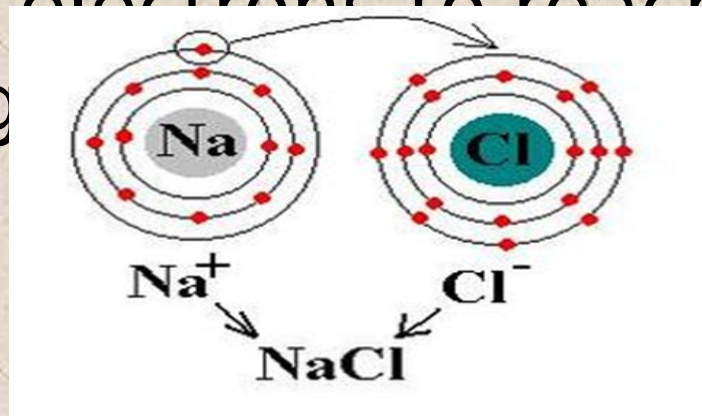
Group	Ion Formation	<i>Electron Configuration</i>	
		Atom	Ion
1	$\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$	$[\text{Ne}]3s^1$ 	$[\text{Ne}]$
2	$\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$	$[\text{Ne}]3s^2$ 	$[\text{Ne}]$
3	$\text{Al} \rightarrow \text{Al}^{3+} + 3\text{e}^-$	$[\text{Ne}]3s^23p^1$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

Section 6.4

Stable Electron Configurations and Charges on Ions

Electron Configurations of Ions

-  Atoms in stable compounds usually have a noble gas electron configuration.
- Metals lose electrons to reach noble gas configuration.
 - Nonmetals gain electrons to reach noble gas configuration.



Section 6.4

Stable Electron Configurations and Charges on Ions

Electron Configurations and Bonding

1. When a *nonmetal* and a *Group 1, 2, or 3 metal* react to form a binary ionic compound, the ions form so that the valence-electron configuration of the nonmetal achieves the electron configuration of the next noble gas atom.
 - **The valence orbital of the metal are emptied to achieve the configuration of the previous noble gas.**

Section 6.4

Stable Electron Configurations and Charges on Ions

2. When *two nonmetals* react to form a covalent bond, they share electrons in a way that completes the valence-electron configurations of both atoms.

Section 6.4

Stable Electron Configurations and Charges on Ions

Predicting Formulas of Ionic Compounds

- Chemical compounds are always electrically neutral.

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]

Section 6.4

Stable Electron Configurations and Charges on



Concept Check

What is the correct electron configuration for the most stable form of the **sulfur ion** in an ionic compound?

- a) $1s^2 2s^2 2p^6 3s^2$
- b) $1s^2 2s^2 2p^6 3s^2 3p^2$
- c) $1s^2 2s^2 2p^6 3s^2 3p^4$
- d) $1s^2 2s^2 2p^6 3s^2 3p^6$

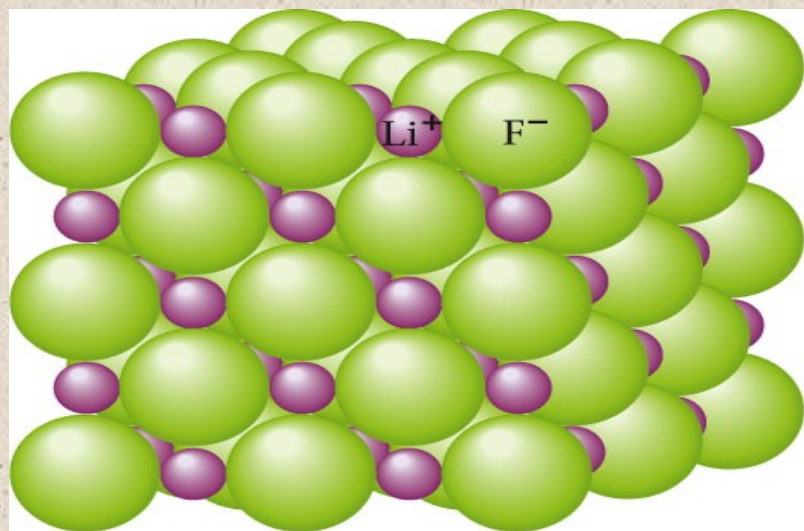
Section 6.5

Forming and Structures of Ionic

Compounds

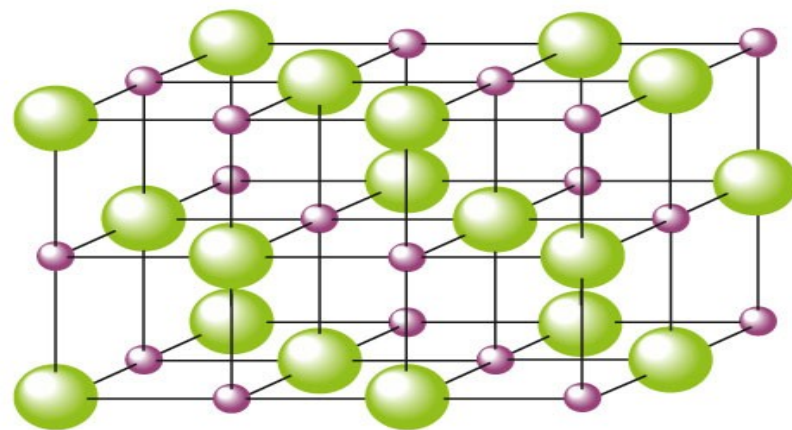
Structures of Ionic Compounds

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



a

This structure represents the ions as packed spheres.



b

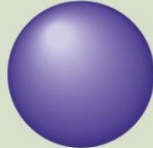
This structure shows the positions (centers) of the ions. The spherical ions are packed in the way that maximizes the ionic attractions.

Section 6.5

Forming and Structures of Ionic Compounds

Structures of Ionic Compounds

- **Cations are always smaller than the parent atom.**
- **Anions are always**

Atom				Cation				Atom				Anion			
Li	152						Li ⁺ 60	F	72				F ⁻ 136		
Na	186						Na ⁺ 95	Cl	99				Cl ⁻ 181		
K	227						K ⁺ 133	Br	114				Br ⁻ 195		
Rb	248						Rb ⁺ 148	I	133				I ⁻ 216		
Cs	265						Cs ⁺ 169								

Section 6.5

Ionic Bonding and Structures of Ionic

Compounds

Concept Check

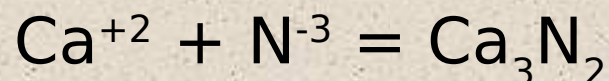
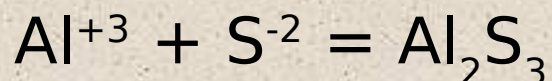
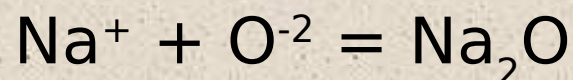
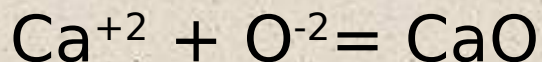
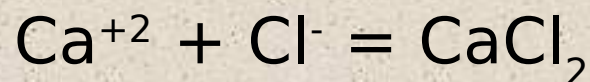
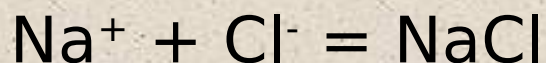


Which atom or ion has the **smallest** radius?

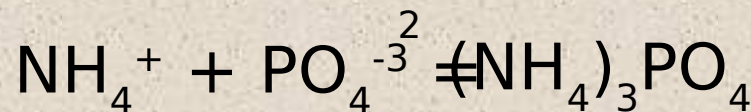
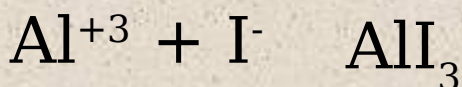
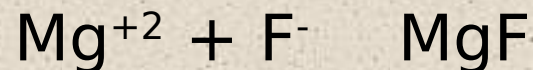
- a) ☐ O^{2+}
- b) ☐ O^{+}
- c) ☐ O
- d) ☐ O^{2-}

Section 6.5

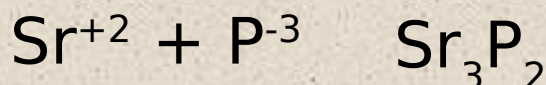
Putting Ions Together



You try these!



=



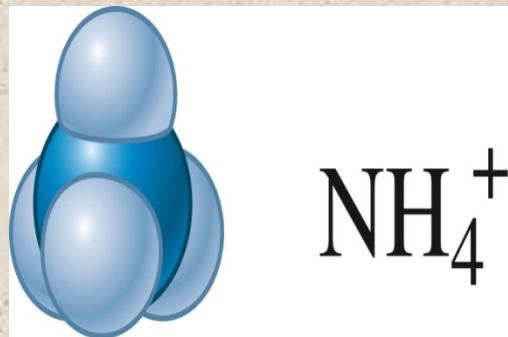
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Section 6.5

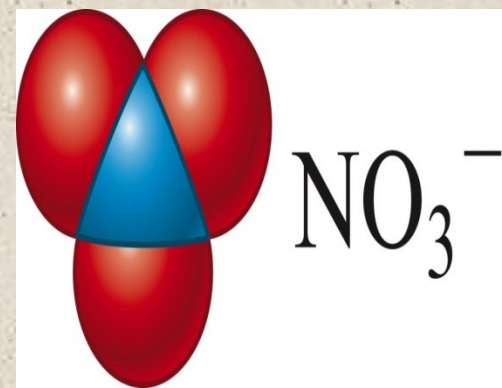
Ionic Bonding and Structures of Ionic Compounds

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.



cation	anion	compound
Ca^{+2}	Cl^{-1}	CaCl_2
Ba^{+2}	O^{-2}	BaO
K^{+1}	S^{-2}	K_2S
Fe^{+3}	Br^{-1}	FeBr_3
Cr^{+3}	O^{-2}	Cr_2O_3

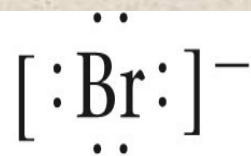


Lewis Structure

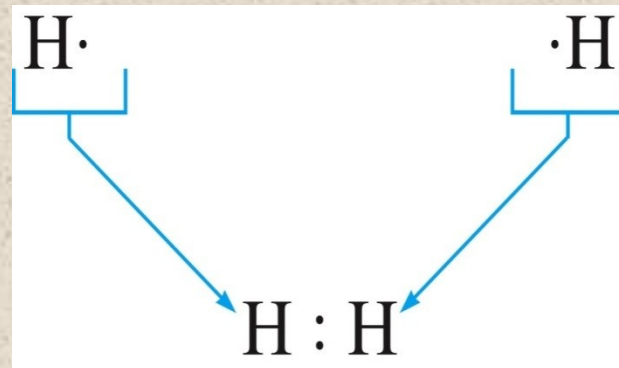
- Shows how valence electrons are arranged among atoms in a molecule.
- Most important requirement
 - Atoms achieve noble gas electron configuration (octet rule, duet rule).



Noble gas
configuration [Ar]

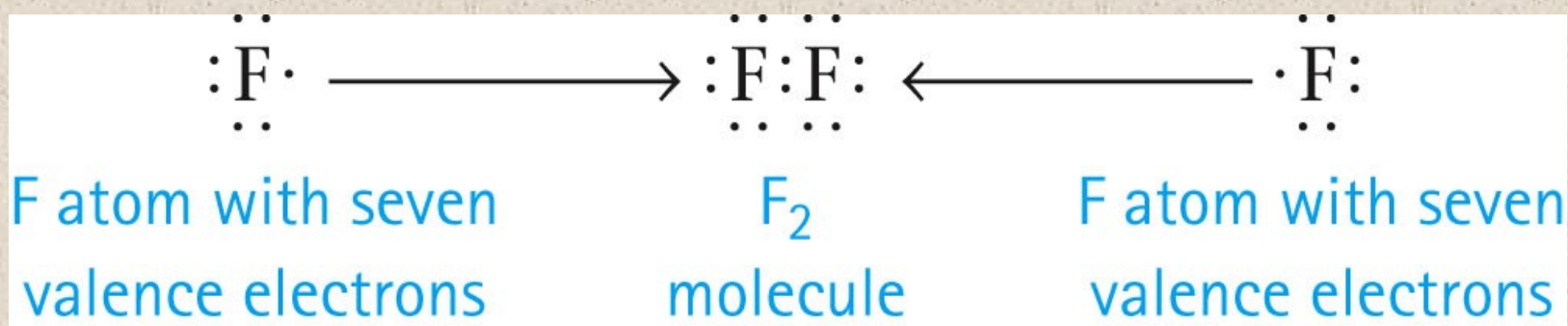


Noble gas
configuration [Kr]



Writing Lewis Structures

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



Section

6.6

1. Write a skeleton molecule with the lone atom in the middle (Hydrogen can never be in the middle)

Rules to Write Dot Structures

**2. Find the number of electrons needed (N)
(8 x number of atoms, 2 x number of H atoms)**

**3. Find the number of electrons you have
(valence e⁻'s) (H)**

**4. Subtract to find the number of bonding
electrons (N-H=B)**

**5. Subtract again to find the number of non-
bonding electrons (H-B=NB)**

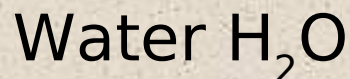
**6. Insert minimum number of bonding electrons
in the skeleton between atoms only. Add more
bonding if needed until you have B bonding
electrons.**

**7. Insert needed non-bonding electrons around
(not between) atoms so that all atoms have 8**

Section 6.6

Lewis Structures

Let's Try it!



1.S $2 \times 2 = 4$ for Hydrogen

$1 \times 8 = 8$ for Oxygen

2.N $4 + 8 = 12$ needed

electrons
 $2 \times 1 = 2$ for Hydrogen

3.H $1 \times 6 = 6$ for Oxygen

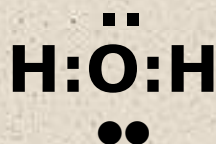
You have 8 available electrons

4.B $12 - 8 = 4$ bonding electrons

5.NB

$8 - 4 = 4$ non-bonding
electrons

6.E



12 N

- 8 H

- 4 B

4 NB



Section 6.6

Lewis Structures

Let's Try it!



Ammonia NH_3

1.S $3 \times 2 = 6$ for Hydrogen

2.N $1 \times 8 = 8$ for Nitrogen

$6 + 8 = 14$ needed

3.H $3 \times 1 = 3$ for Hydrogen

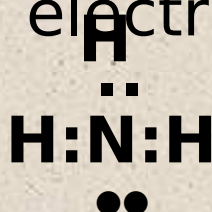
$1 \times 5 = 5$ for Nitrogen

You have 8 available electrons

4.B $14 - 8 = 6$ bonding electrons

5.NB $8 - 6 = 2$ non-bonding

6.E electrons

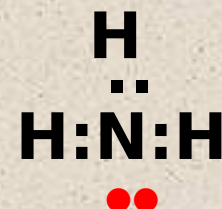
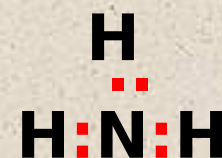


14 N

- 8 H

- 6 B

2 NB



Section 6.6

Lewis Structures

Let's Try it!



Carbon Dioxide



1.S

1 x 8 = 8 for Carbon

2.N

2 x 8 = 16 for Oxygen

8+16=24 needed

24 N

3.H

~~1 x 4 = 4 for Carbon~~

2 x 6 = 12 for Oxygen

You have 16 available electrons

- 16 H

- 8 B

4.B

24 - 16 = 8 bonding electrons

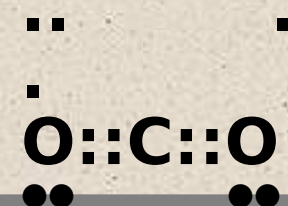
8 NB

5.NB

16 - 8 = 8 non-bonding electrons



6.E



Section 6.6

Lewis Structures

Let's Try it!



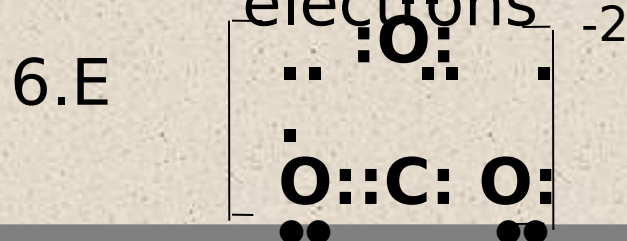
Carbonate CO₃⁻²

2.N 3 x 8 = 24 for Oxygen
 1 x 8 = 8 for Carbon
 24+8=32 needed electrons

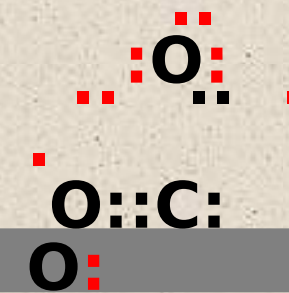
3.H 3 x 6 = 18 for Oxygen
 1 x 4 = 4 for Carbon
 You have 22 + 2 more
 available e-'s

4.B 32 - 24 = 8 bonding electrons

5.NB 24 - 8 = 16 non-bonding
 electrons



$$\begin{array}{r} 32 \text{ N} \\ - 24 \text{ H} \\ \hline 8 \text{ B} \\ \hline 16 \text{ NB} \end{array}$$



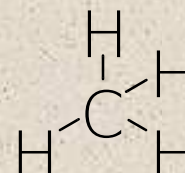
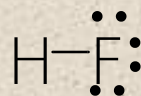
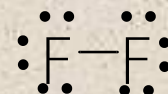
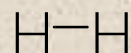
Section 6.6

Lewis Structures



Concept Check

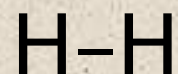
Draw a Lewis structure for each of the following molecules:



Section 6.7

Lewis Structures of Molecules with Multiple Bonds

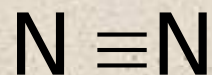
- **Single bond** – covalent bond in which 1 pair of electrons is shared by 2 atoms.



- **Double bond** – covalent bond in which 2 pairs of electrons are shared by 2 atoms.



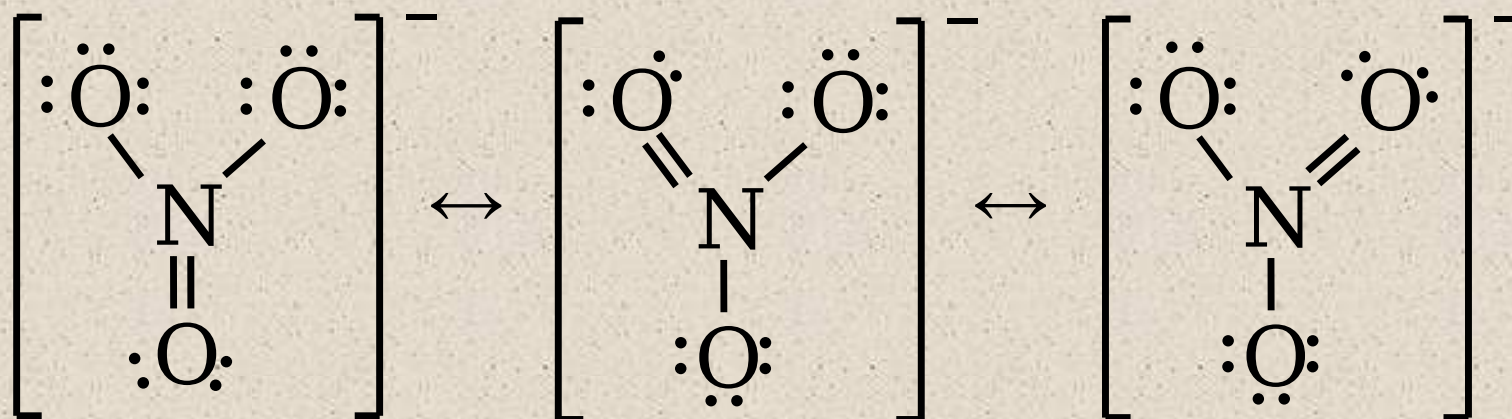
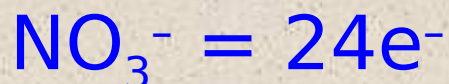
- **Triple bond** – covalent bond in which 3 pairs of electrons are shared by 2 atoms.



Section 6.7

Lewis Structures of Molecules with Multiple Bonds Resonance

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

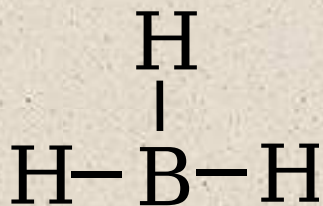
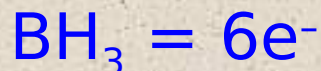


Section 6.7

Lewis Structures of Molecules with Multiple Bonds

Some Exceptions to the Octet Rule

- Boron tends to form compounds in which the boron atom has fewer than eight electrons around it (it does not have a complete octet).



- Molecules containing odd numbers of electrons like NO and NO₂.

Section 6.7

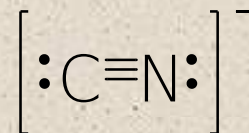
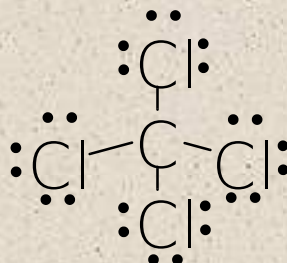
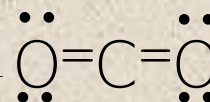
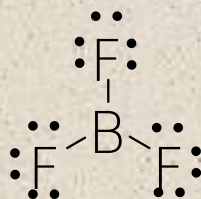
Lewis Structures of Molecules with Multiple



ds

Concept Check

Draw a Lewis structure for each of the following molecules:



Section 6.7

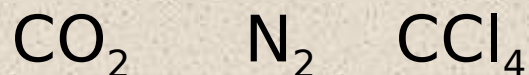
Lewis Structures of Molecules with Multiple



ds

Concept Check

Consider the following compounds:

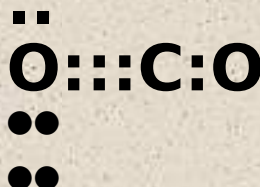


Which compound exhibits **resonance**?

a)



b)



c)



d)

At least two of the above compounds exhibit resonance.

Section 6.7

Lewis Structures of Molecules with Multiple



ds

Concept Check

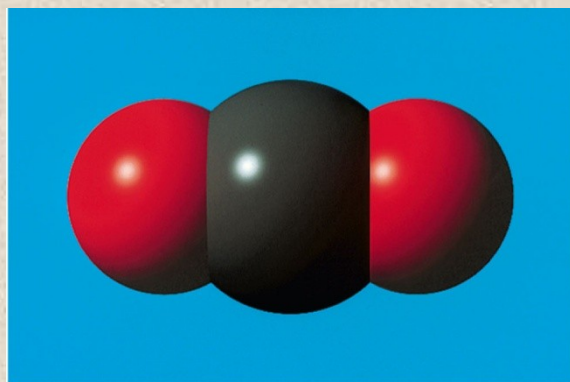
Which of the following supports why Lewis structures are **not** a completely accurate way to draw molecules?

- a) We cannot say for certain where an electron is located yet when drawing Lewis structures, we assume the electrons are right where we place them.
- b) When adding up the number of valence electrons for a molecule, it is possible to get an odd number which would make it impossible to satisfy the octet rule for all atoms.
- c) Both statements 1 and 2 above support why Lewis structures are not a completely accurate way to draw molecules.
- d) Lewis structures are the most accurate way to draw molecules and are completely correct.

Section 6.8

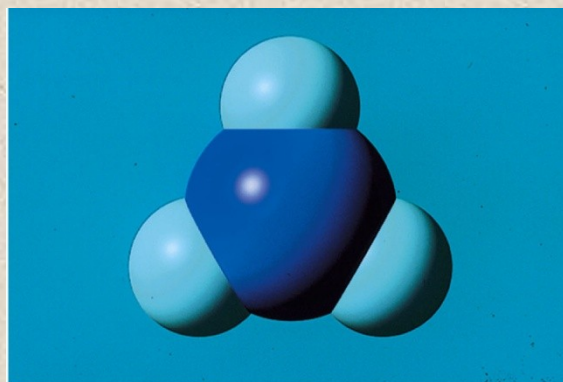
Molecular Structure

- Three dimensional arrangement of the atoms in a molecule.



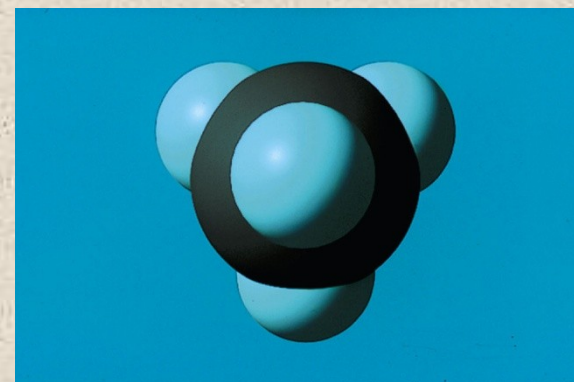
a

Computer graphic of a linear molecule containing three atoms



b

Computer graphic of a trigonal planar molecule



c

Computer graphic of a tetrahedral molecule

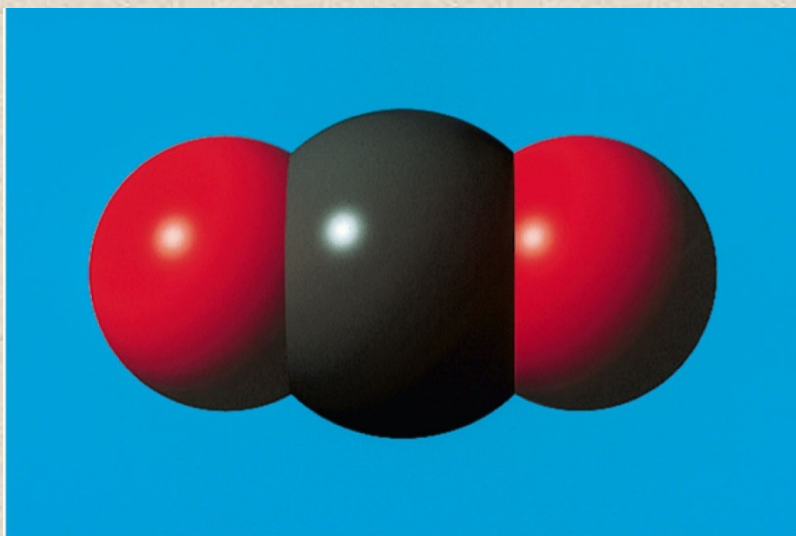
Bond Angle 180°
 $\sim 109^\circ$

$\sim 120^\circ$

Section 6.8

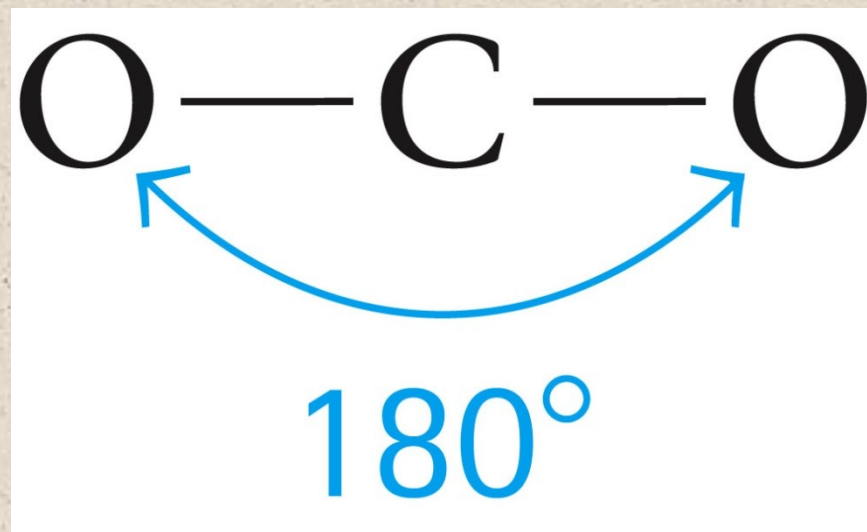
Molecular Structure

- Linear structure – atoms in a line
 - Carbon dioxide



a

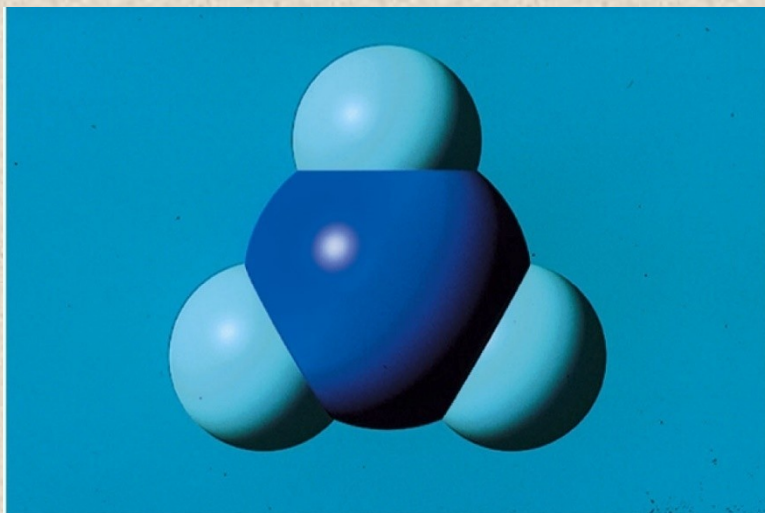
Computer graphic of a linear molecule containing three atoms



Section 6.8

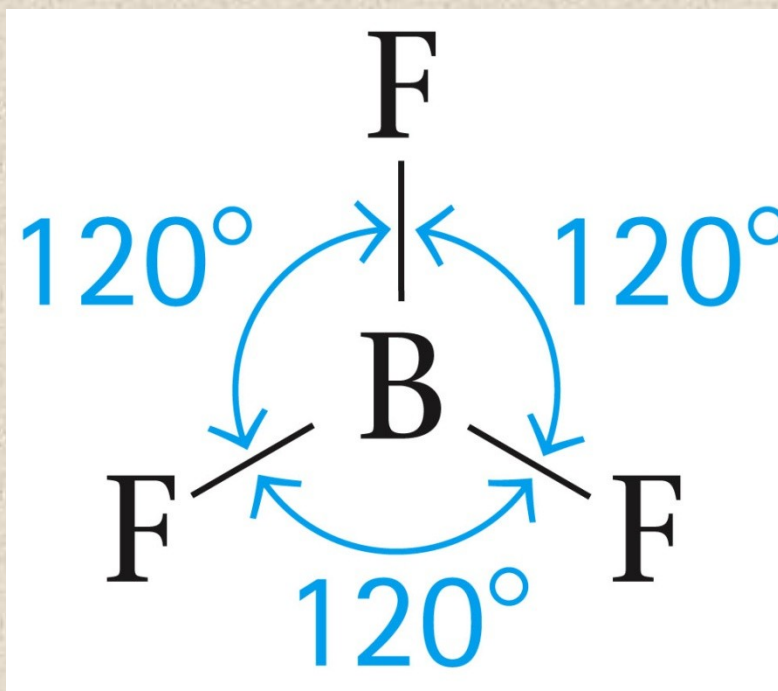
Molecular Structure

- Trigonal planar – atoms in a triangle
 - Boron trifluoride



b

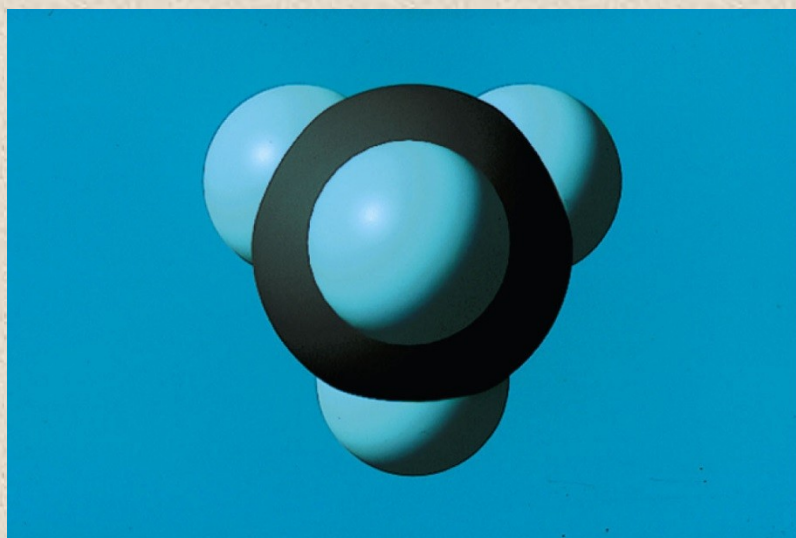
Computer graphic of a trigonal planar molecule



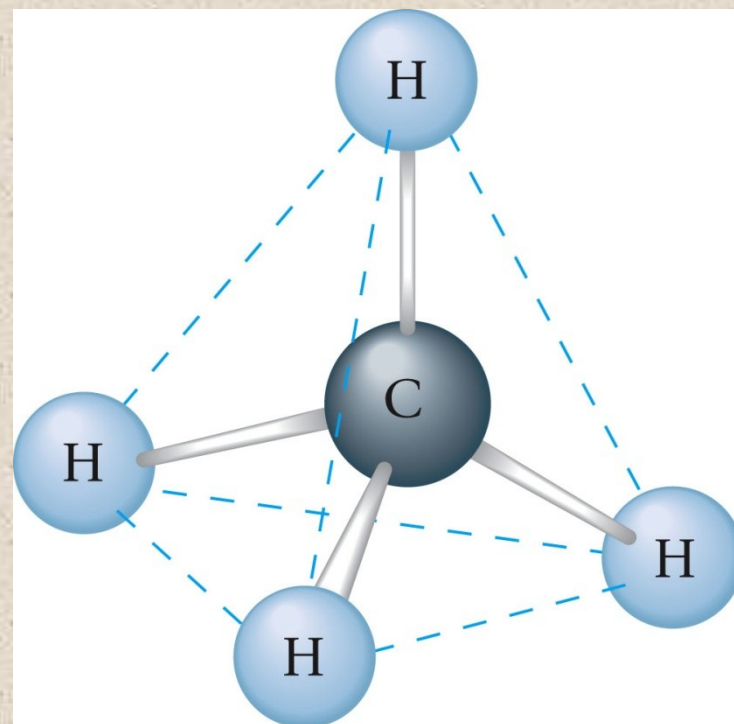
Section 6.8

Molecular Structure

- Tetrahedral structure
 - Methane



Computer graphic of a tetrahedral molecule



Section 6.8

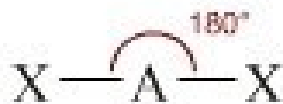
Molecular Structure

Molecular Shapes and bond angles from Lab Book Figure

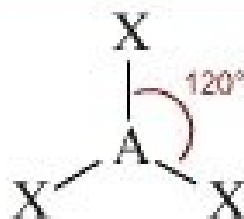
2 Electron Grps = 180°

3 Electron Grps = $\sim 120^\circ$

4 Electron Grps = $\sim 109^\circ$



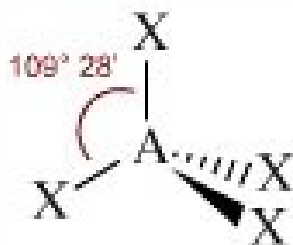
AX_2
Linear



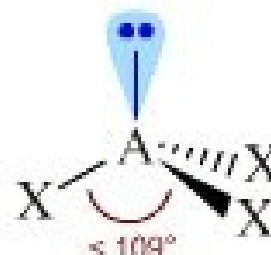
AX_3
Trigonal planar



AX_2E_1
Bent or Angular



AX_4
Tetrahedral



AX_3E_1
Trigonal pyramidal

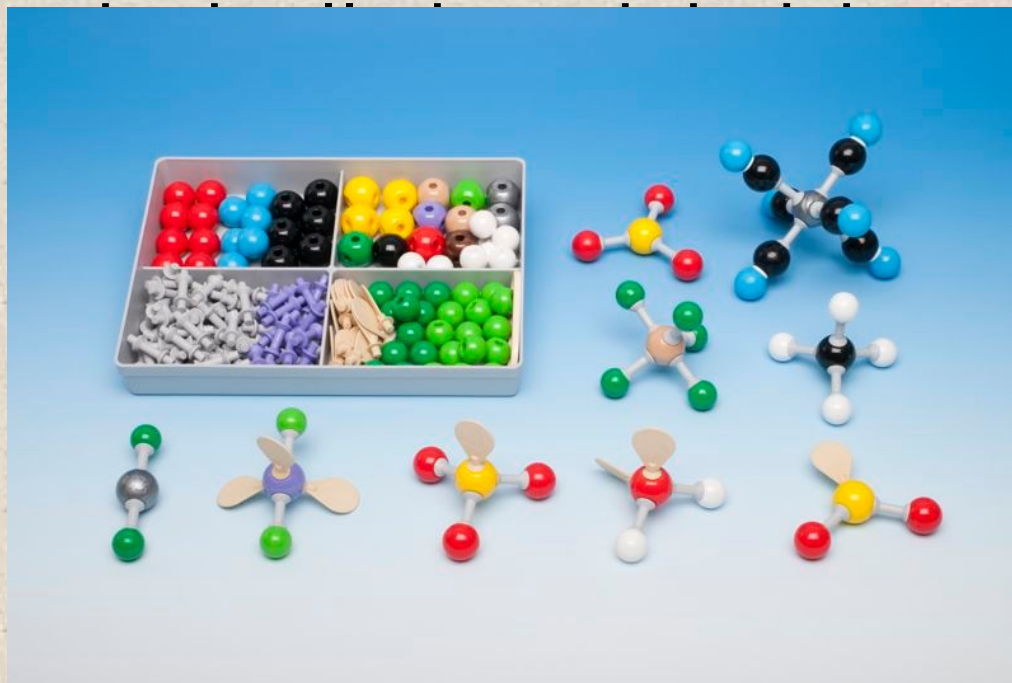


AX_2E_2
Bent or Angular

6.9 Molecular Structure

VSEPR Model

- VSEPR: Valence Shell Electron-Pair Repulsion.
- The structure around a given atom is determined by the number of valence electron pairs



Section 6.9

Molecular Structure

Two Pairs of Electrons

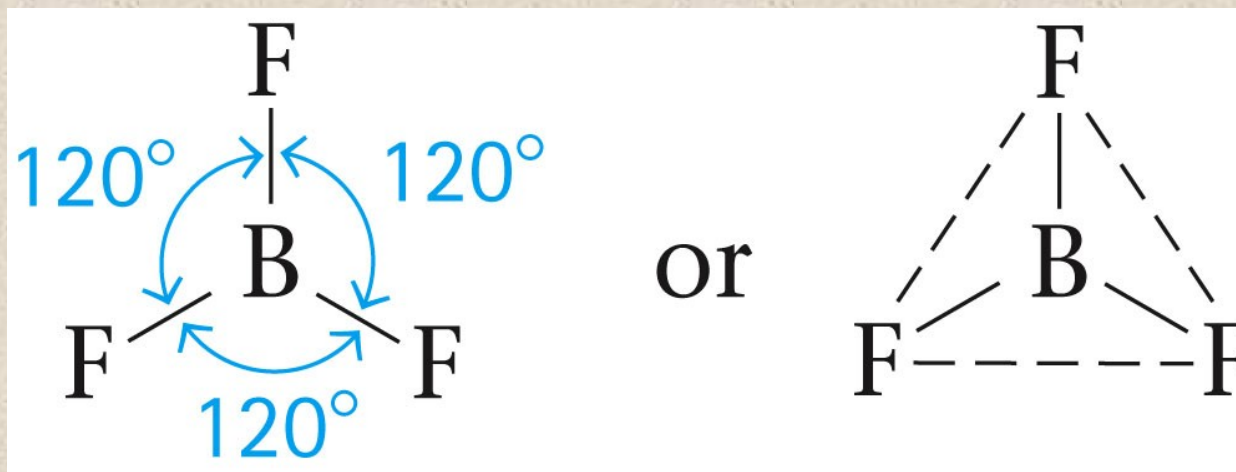
- BeCl_2
- 180°
- Linear



Molecular Structure

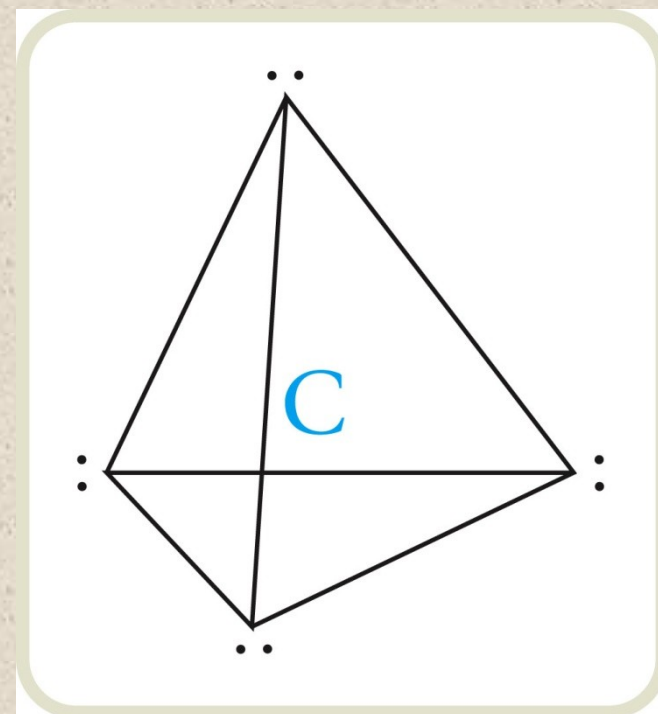
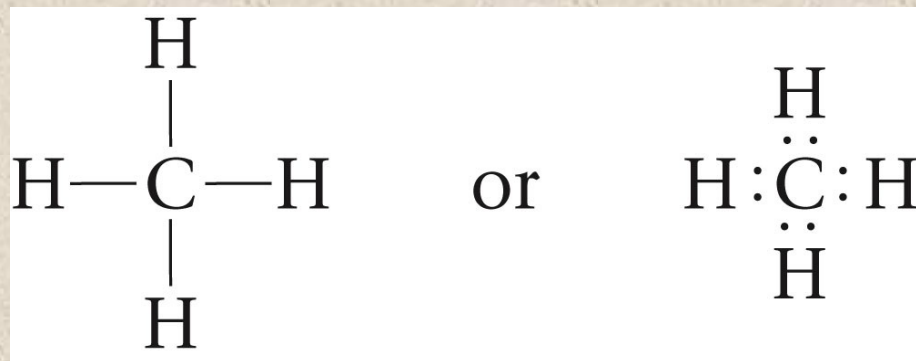
Three Pairs of Electrons

- BF_3
- 120°
- Trigonal planar



Four Pairs of Electrons

- CH_4
- 109.5°
- Tetrahedral



Molecular Structure

Steps for Predicting Molecular Structure Using the VSEPR Model

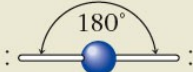
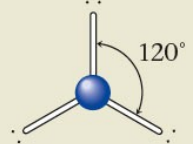
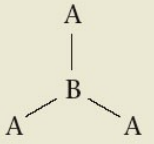
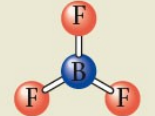
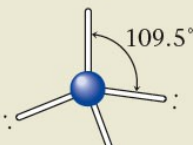
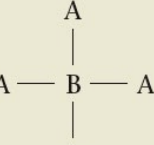
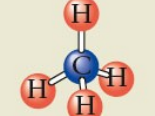
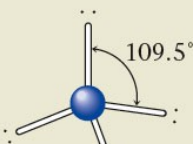
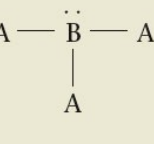
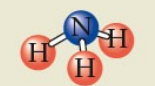
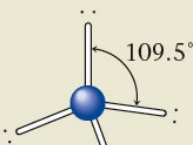
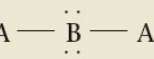
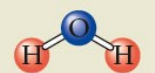
1. Draw the Lewis structure for the molecule.
2. Count the electron pairs and arrange them in the way that minimizes repulsion (put the pairs as far apart as possible).
3. Determine the positions of the atoms from the way electron pairs are shared (how electrons are shared between the central atom and surrounding atoms).
4. Determine the name of the molecular structure from positions of the atoms.

Section 6.9

Molecular Structure

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

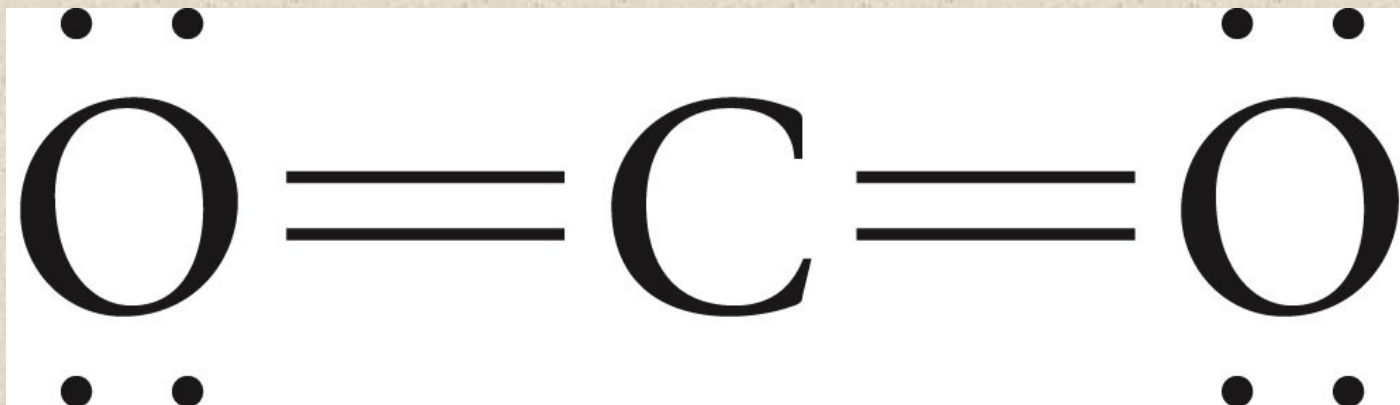
Table 12.4 Arrangements of Electron Pairs and the Resulting Molecular Structures for Two, Three, and Four Electron Pairs

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		

Molecular Structure

Molecules with Double Bonds

- When using the VSEPR model to predict the molecular geometry of a molecule, a double or triple bond is counted the same as a single electron pair.



Section 6.10

L0

Molecular Structure



Concept Check

Determine the **molecular structure** for each of the following molecules, and include **bond angles**:



HCN – linear, 180°

PH₃ – trigonal pyramid,

$109.5^\circ (107^\circ)$

SeO₂ – bent, 120°