Bonding

Ionic bonding – the bonds that hold ionic compounds together

- start w/ neutral atoms — $K_{(g)}$ atom and $Br_{(g)}$ atom
- transfer of valence electrons — $K_{(g)}$ becomes $K^+_{(g)}$ $Br_{(g)}$ atom becomes $Br^-_{(g)}$
- electrostatic attraction between ions — $K^+_{(g)} + Br^-_{(g)} \rightarrow KBr$

Covalent bonding – involves electrons being shared
- always valence e⁻

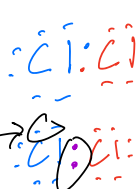
Lewis structures – describe valence e⁻ as dot

atoms want "octet" / noble gas config of e⁻

not

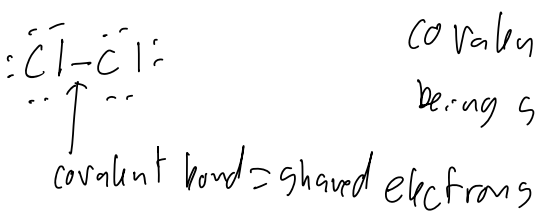
Imagine 2 Cl atoms

bring them near to each other so they can share valence e⁻



each Cl feels it has 8e⁻ – 6 of its own and 2e⁻ that are shared

"lone pairs" →

Cl_2 

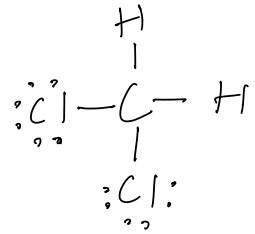
How to make good Lewis structures

1. Find total # of valence e⁻
2. Arrange atoms - determine connectivity
   - central atoms vs peripheral atoms
   - how many bonds atoms tend to form
     - H, F, Cl, Br, I tend to form 1 bond
     - O, S, Se tend to form 2 bonds
     - N, P tend to form 3 bonds
     - C, Si, Ge tend to form 4 bonds
3. Connect all atoms w/ bonds (each bond "uses" 2 valence e⁻)
4. Distribute remaining valence e⁻ as lone pairs to get atoms to their octet
5. Check - # of e⁻, octet
6. If some atoms don't have octet - consider an additional bond for more sharing

* H doesn't want octet, only wants 2 valence e⁻

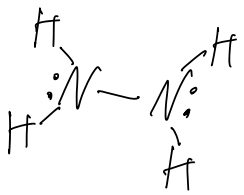
CH_2Cl_2

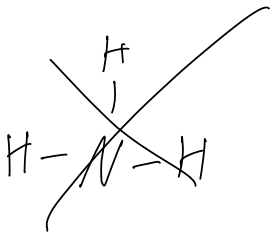
4 2×1 2×7 = 20 e⁻



N_2H_4

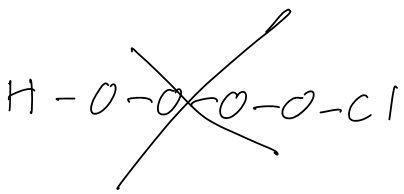
2×5 4×1 = 14 e⁻

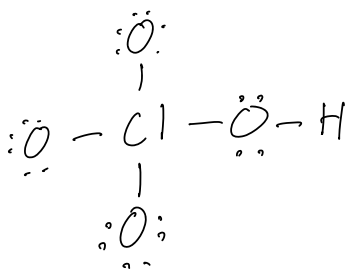


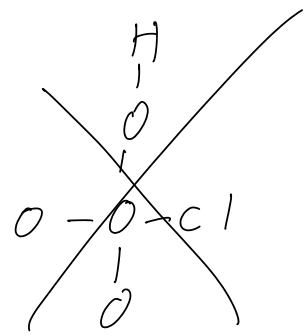


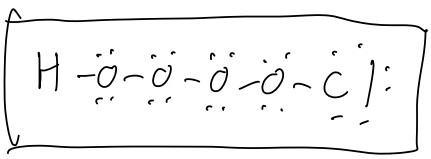
$HClO_4$

1 7 4×6 = 32 e⁻









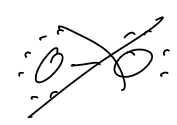
– even though it follows the "rules" it isn't a realistic representation of $HClO_4$

N_2 – 10e

$:N \equiv N:$

└ triple bond

O_2 – 12e⁻



$:\ddot{O}=\ddot{O}:$

└ double bond

CO-love $:C\equiv O:$ CO and N_2 are isoelectronic

Formal charge
1. Bisect all bonds on a molecule and "give" 1e to each atom involved
2. Count all e⁻ (lone pairs + e⁻ from bond chopping) around each atom
3. Compare (2) to # of electrons on neutral atom would have

$:N\equiv N:$

each N will have
2e⁻ from lone pair → 5e⁻
3e⁻ from bond chopping

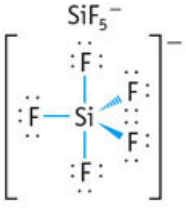
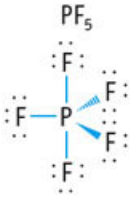
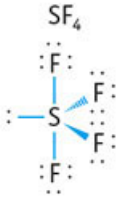
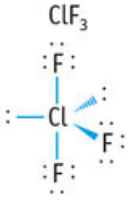
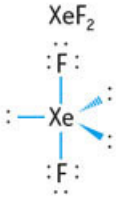
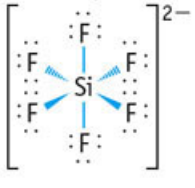
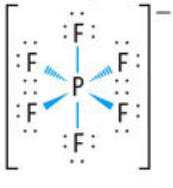
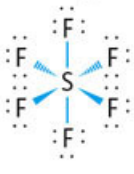
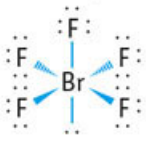
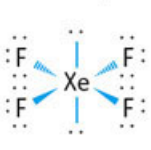
neutral N → 5e⁻

$:N\equiv N:$
FC: 0 FC: 0

$:C\equiv O:$
5e⁻ 5e⁻
FC: -1 FC: +1

when FC ≠ 0 atoms are unhappy → molecules is reactive

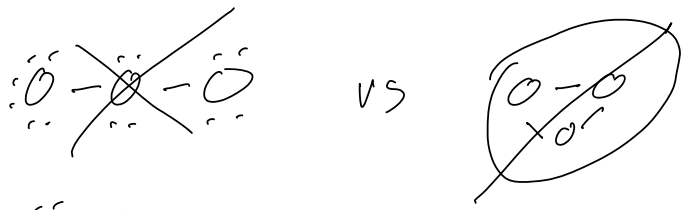
TABLE 8.6 Lewis Structures in Which the Central Atom Exceeds an Octet

Group 4A	Group 5A	Group 6A	Group 7A	Group 8
SiF_5^- 	PF_5 	SF_4 	ClF_3 	XeF_2 
SiF_6^{2-} 	PF_6^- 	SF_6 	BrF_5 	XeF_4 

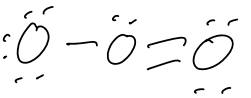
H is happy w/ 2e

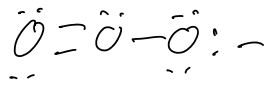
B is happy w/ 6e

O_3 – 18e⁻



vs

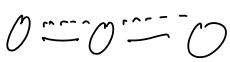
 – good structure

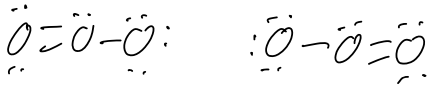
 – same structure

This structure suggests that the two sides of molecule are different

But experiments show that molecule is symmetric

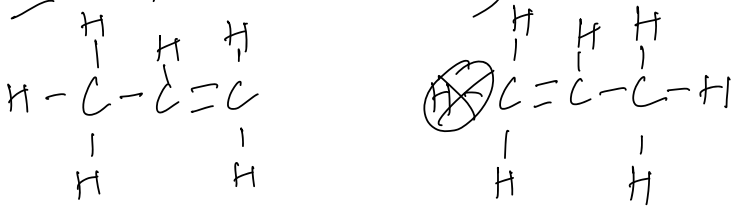
"resonance" – the 2 electrons shown in double bond are really being shared by all 3 atoms





referred to as "resonance structures" of O_3

*For resonance to happen the double bond has to be able to be in more than one place without any atoms moving



not resonance

Valence shell Electron Pair Repulsion (VSEPR)

describes 3-D structure

* electron pairs want to avoid each other b/c they repel

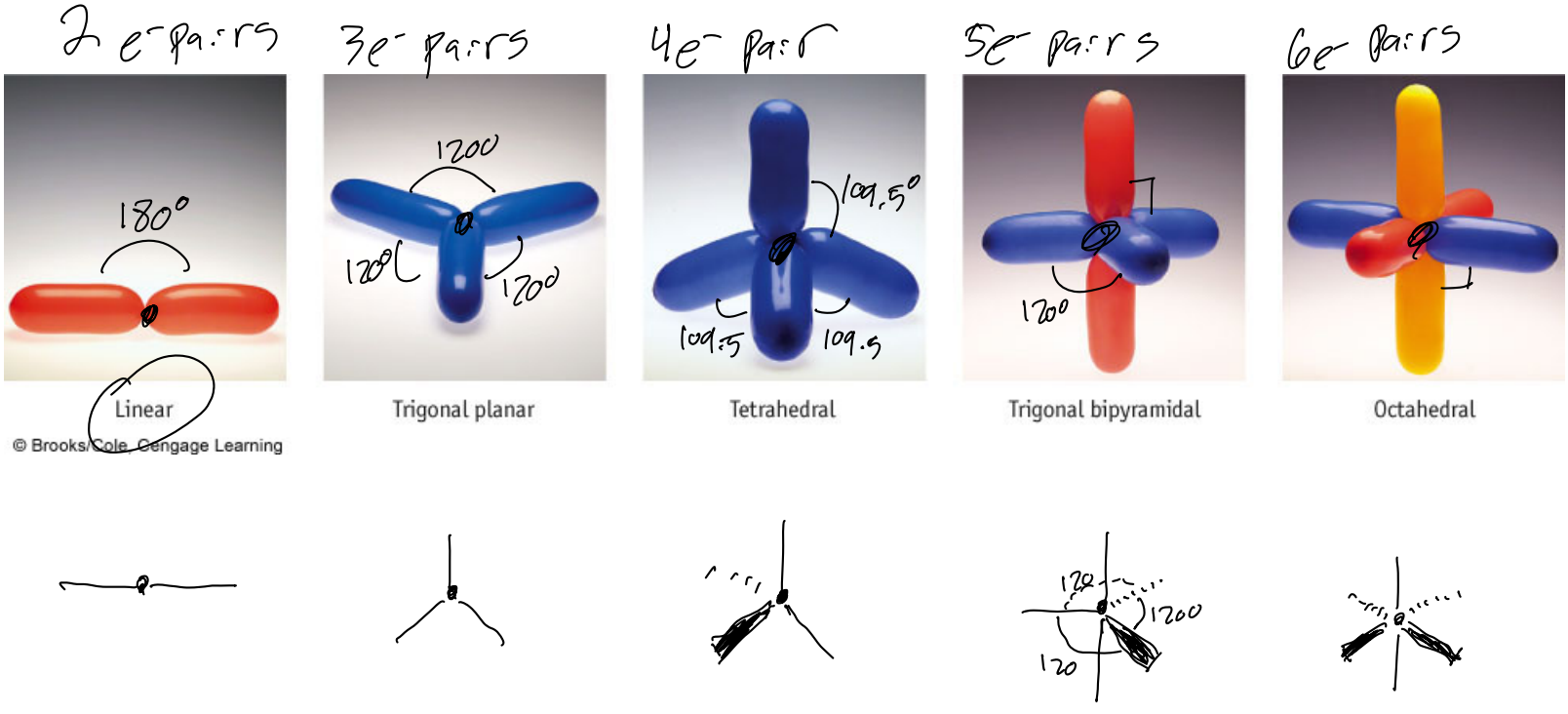

2 e⁻ pairs
3e⁻ pairs
4e⁻ pair
5e⁻ pairs
6e⁻ pairs
180°
120°
120°
120°
109.5°
109.5
109.5
120°
Linear
Trigonal planar
Tetrahedral
Trigonal bipyramidal
Octahedral
© Brooks/Cole, Cengage Learning

Fig. 8-4, p. 368

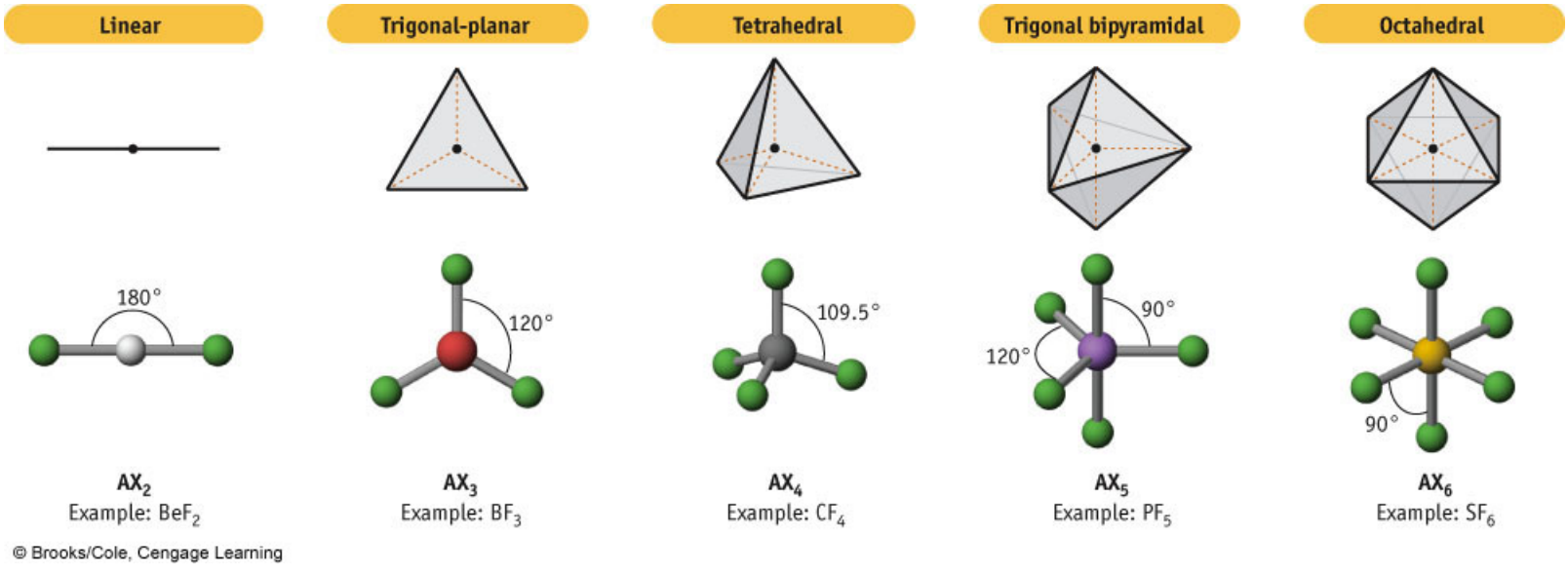


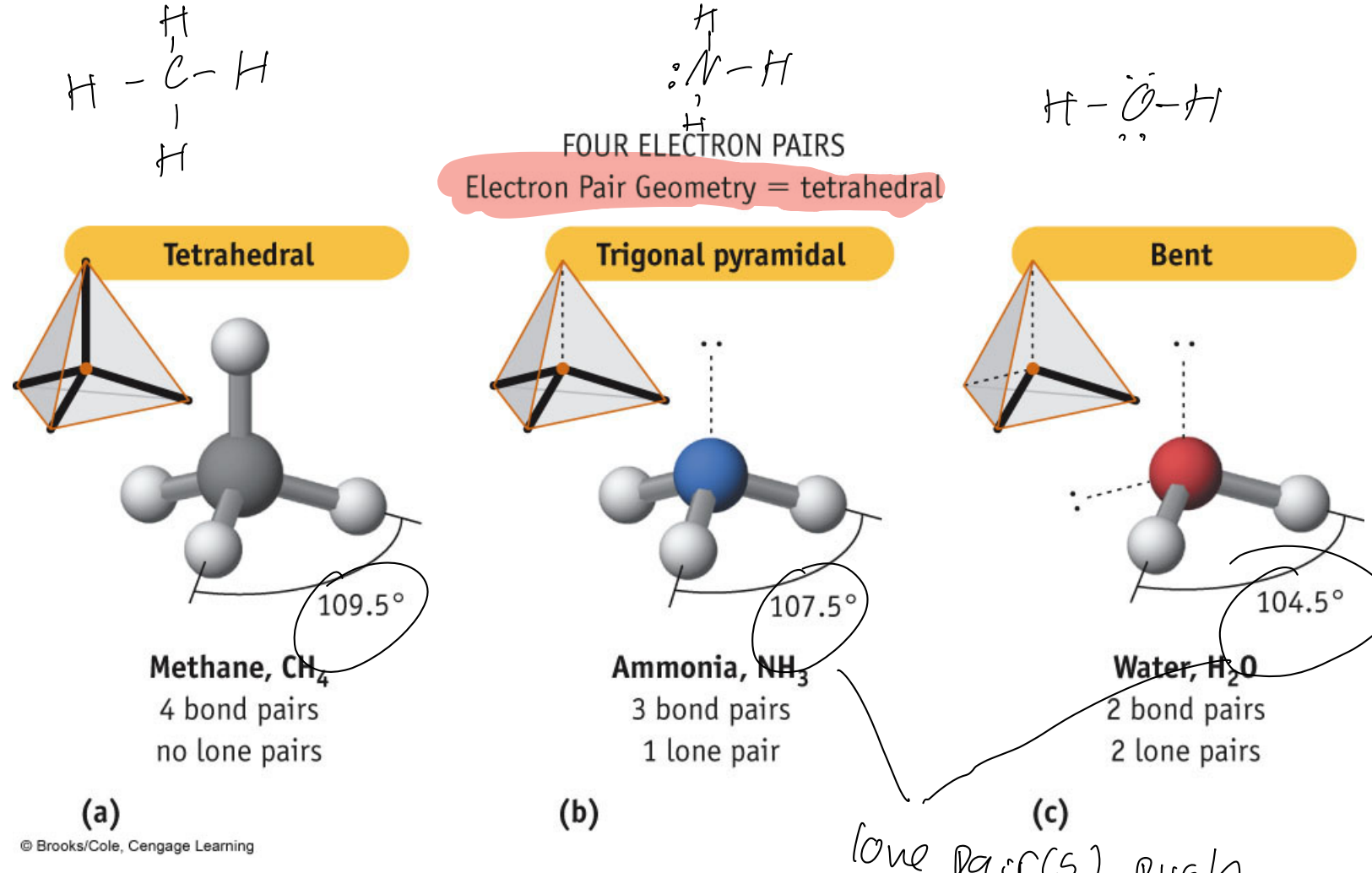
Fig. 8-5, p. 369

consider each atom with 2 or more connected to it

How many electron pairs surround that atom

How many are bonds and how many are lone pairs

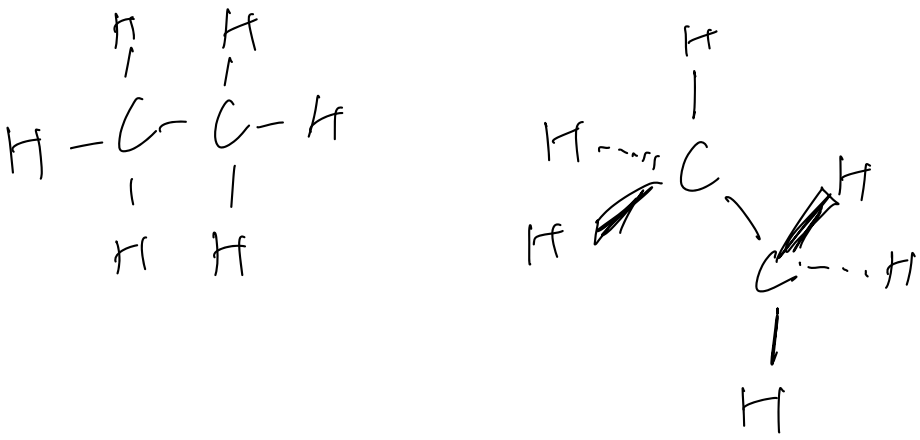
Electron pair geometry doesn't always = molecular geometry



lone pair(s) push bonds to smaller angle b/c repulsion

Molecular geometry = where atoms are

Fig. 8-6, p. 371



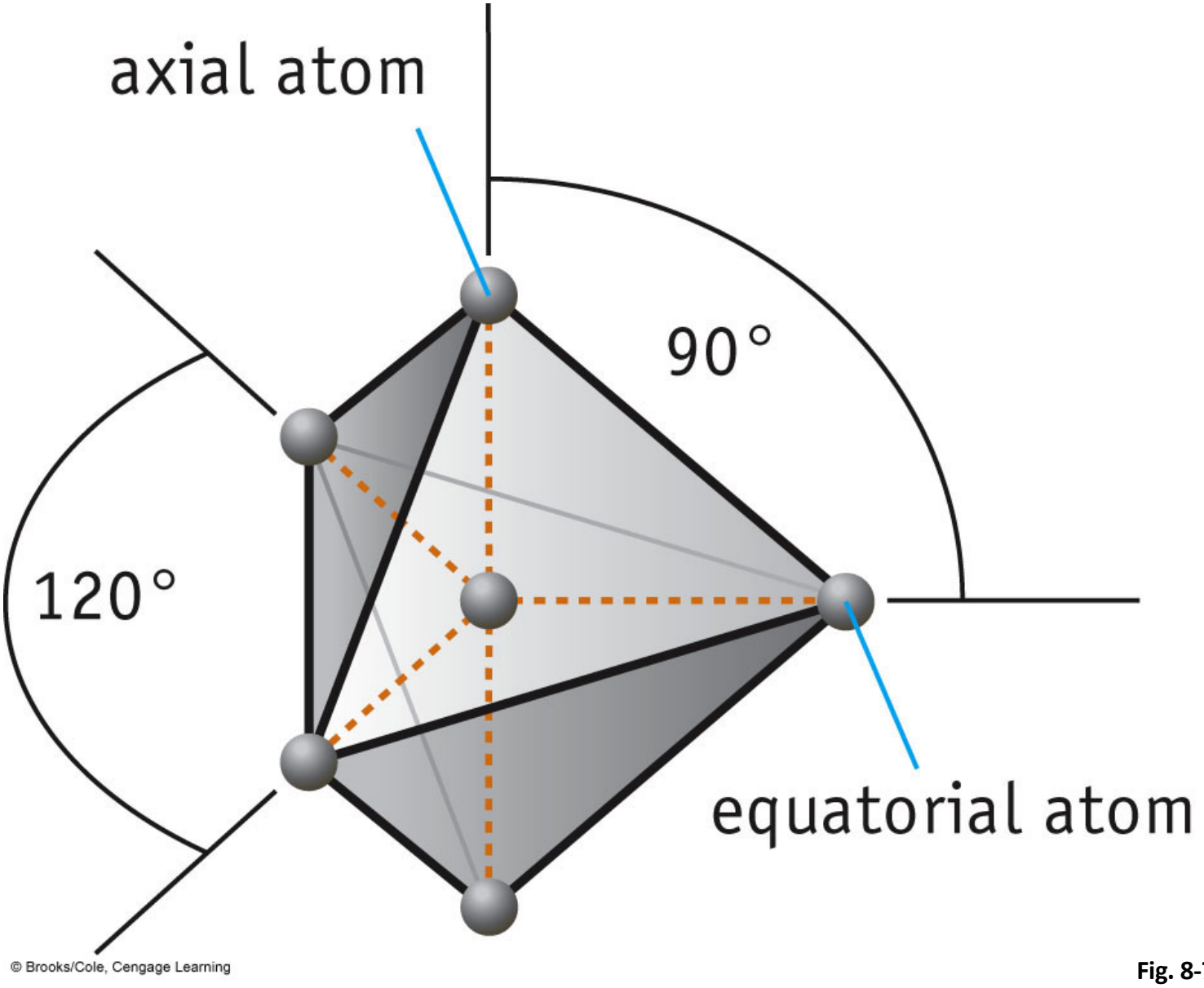
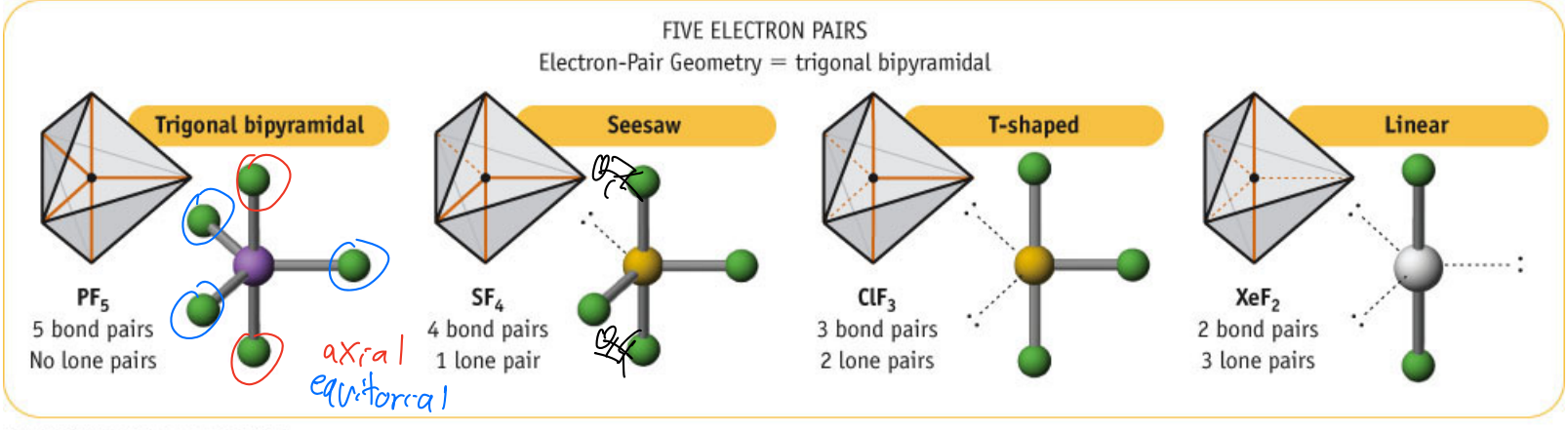
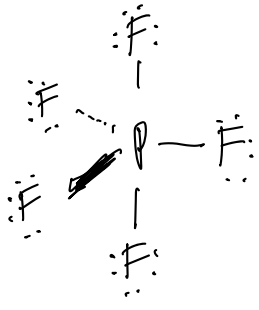
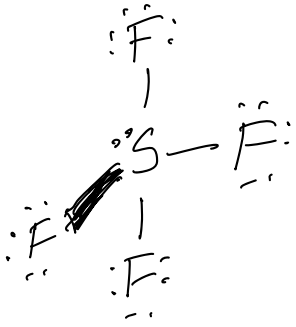


Fig. 8-7, p. 372

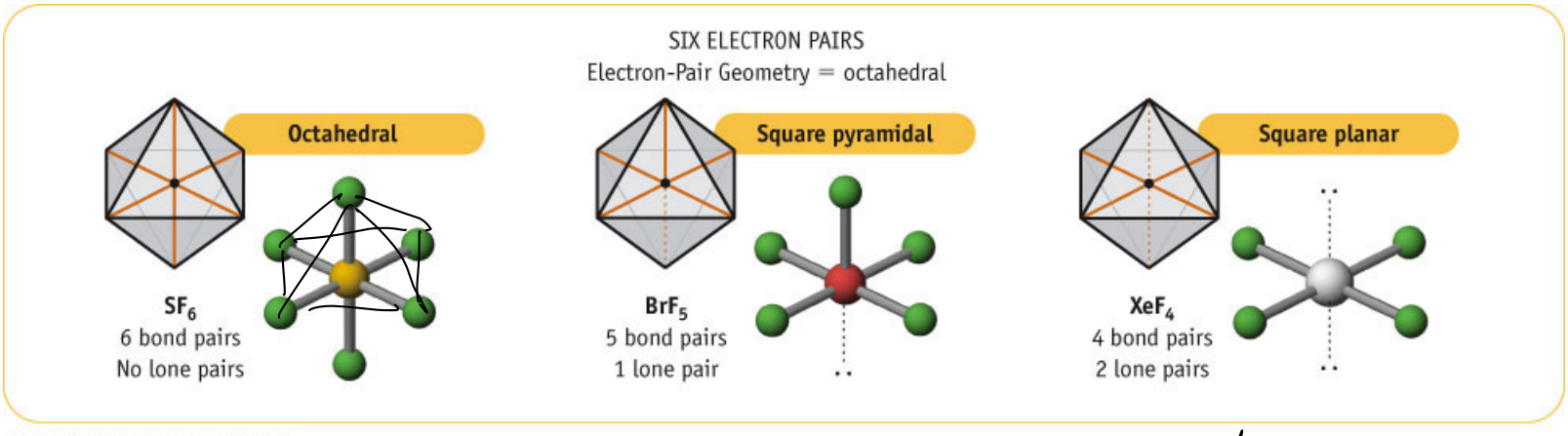




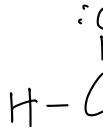


equitorial positions are where lonepairs go

Fig. 8-8, p. 372

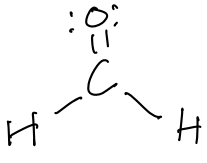


* double/triple bond treated as one electron pair for geometry



- treat as if 3 e-pairs around C

→ trigonal planar



* 2nd bond to lone pair swap removes bond opposite to 1st lone pair

Fig. 8-8, p. 372

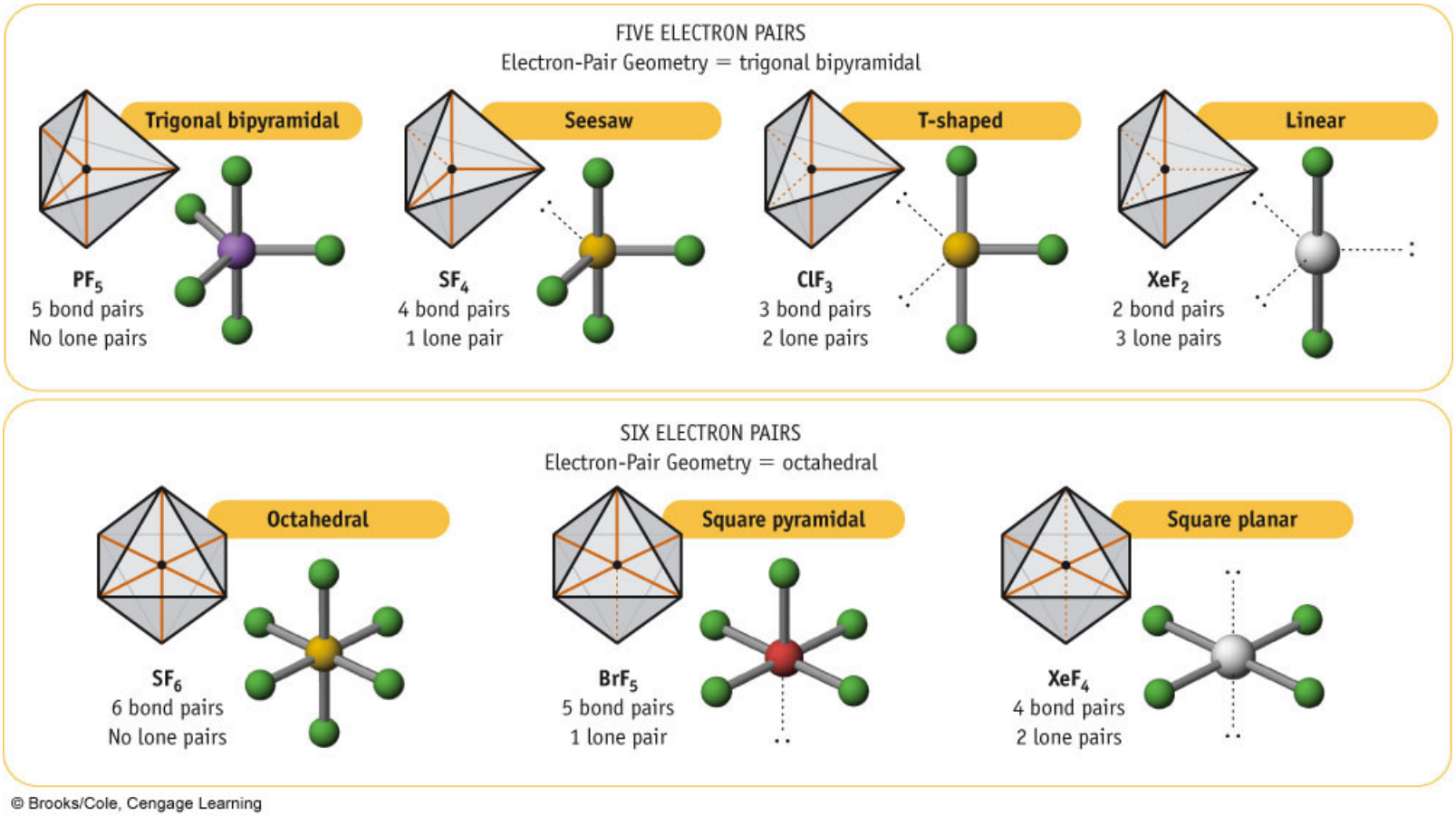


Fig. 8-8, p. 372

Ideal bond has equitable sharing of electrons
often this is not the case

one end of bond becomes more electron rich and other end ~~become~~ becomes more electron poor → a polar bond

Electronegativity is a property of atoms that describes how "selfish" atoms will be with electrons

Difference in electronegativity between atoms determines bond polarity $\Delta > 0.5$ means a polar bond

1A	2A											3A	4A	5A	6A	7A
								H 2.2								
Li 1.0	Be 1.6											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.3	3B	4B	5B	6B	7B	8B			1B	2B	Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2
K 0.8	Ca 1.0	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 1.9	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 1.9	Te 2.1	I 2.7
Cs 0.8	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 2.4	Re 1.9	Os 2.2	Ir 2.2	Pt 2.3	Au 2.5	Hg 2.0	Tl 1.6	Pb 2.3	Bi 2.0	Po 2.0	At 2.2

<1.0 1.0–1.4 1.5–1.9 2.0–2.4 2.5–2.9 3.0–4.0

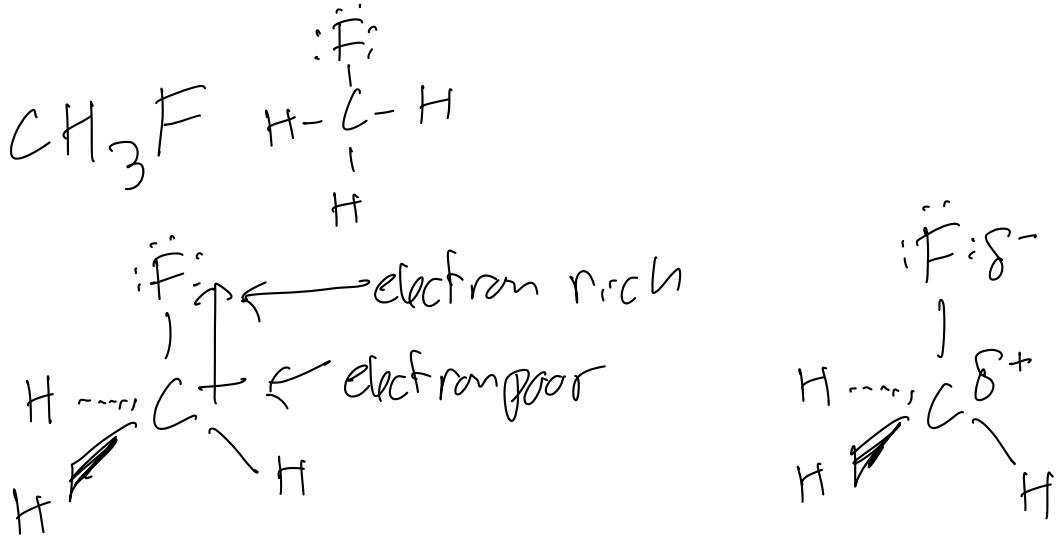
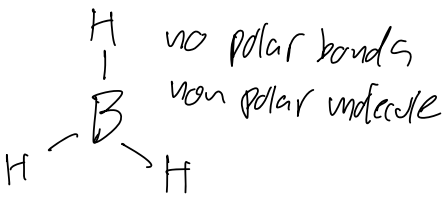


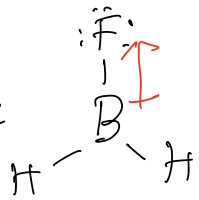
Fig. 8-11, p. 376

* Can also talk about whole molecule being polar or not

BH_3



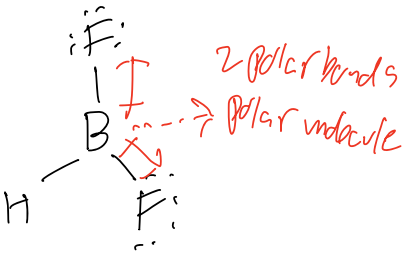
no polar bonds
non polar molecule



1 polar bond
polar molecule

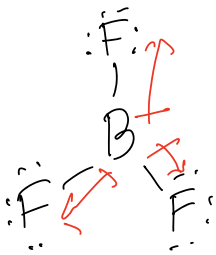
BH_2F

BHF_2



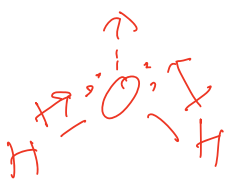
2 polar bonds
polar molecule

BF_3



3 polar bonds
non polar molecule

H_2O



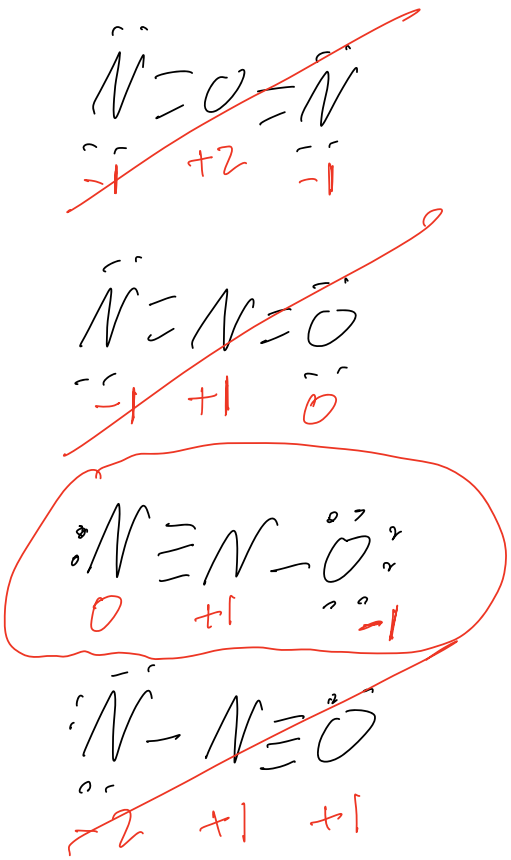
2 polar bonds
polar molecule

One more thing w/ electronegativity

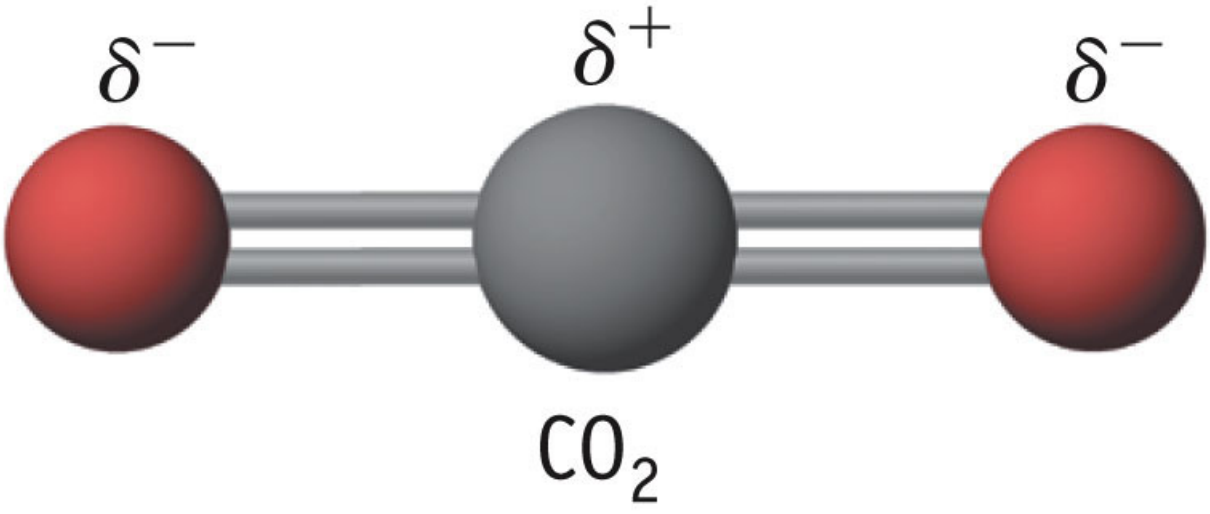
Use w/ lewis structures and formal charge

N_2O - 16e⁻

* more neg formal charge ideally live on more electroneg atoms

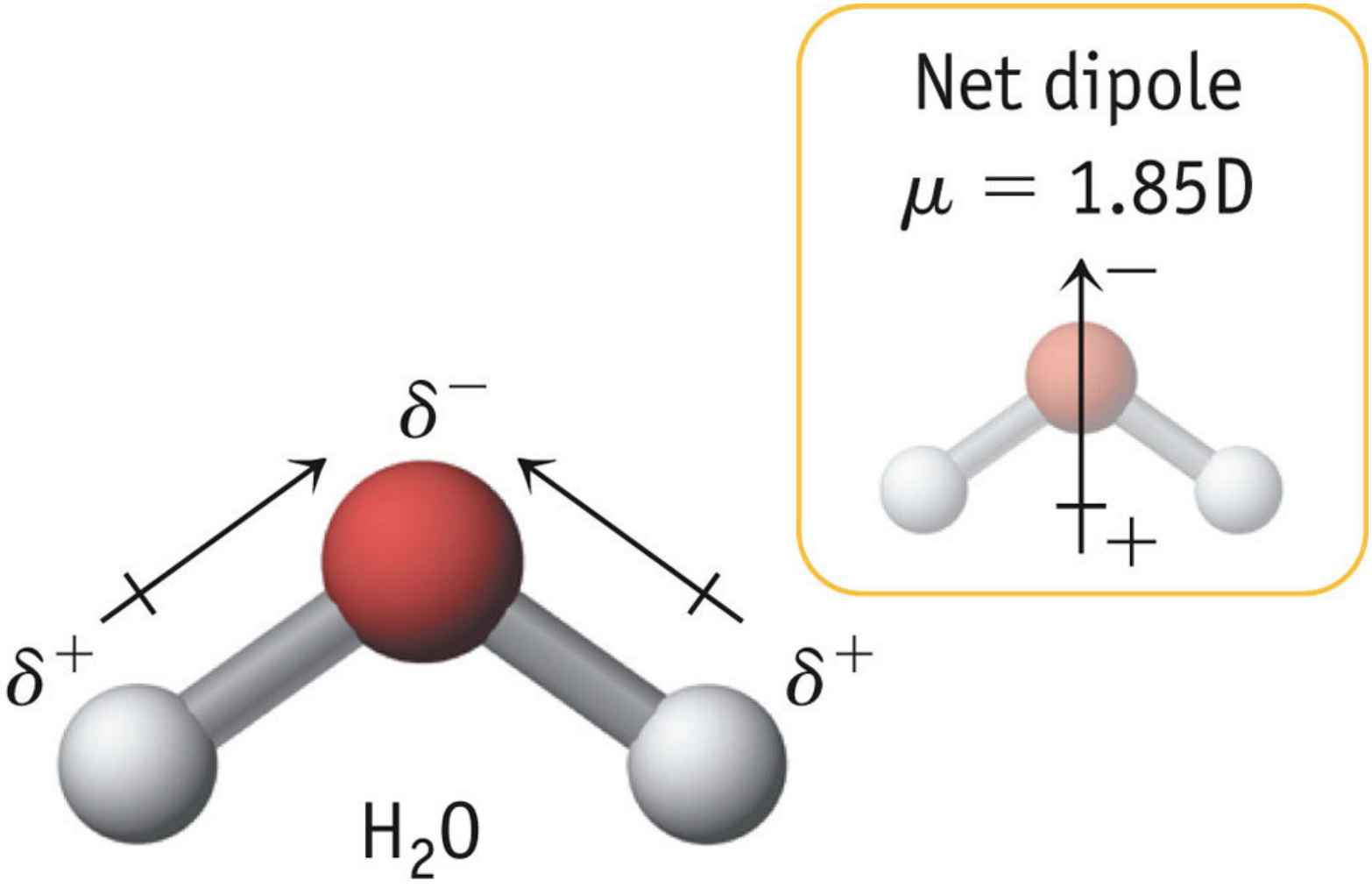


No net dipole moment



(a)

Fig. 8-13, p. 381



(b)

Fig. 8-13, p. 381

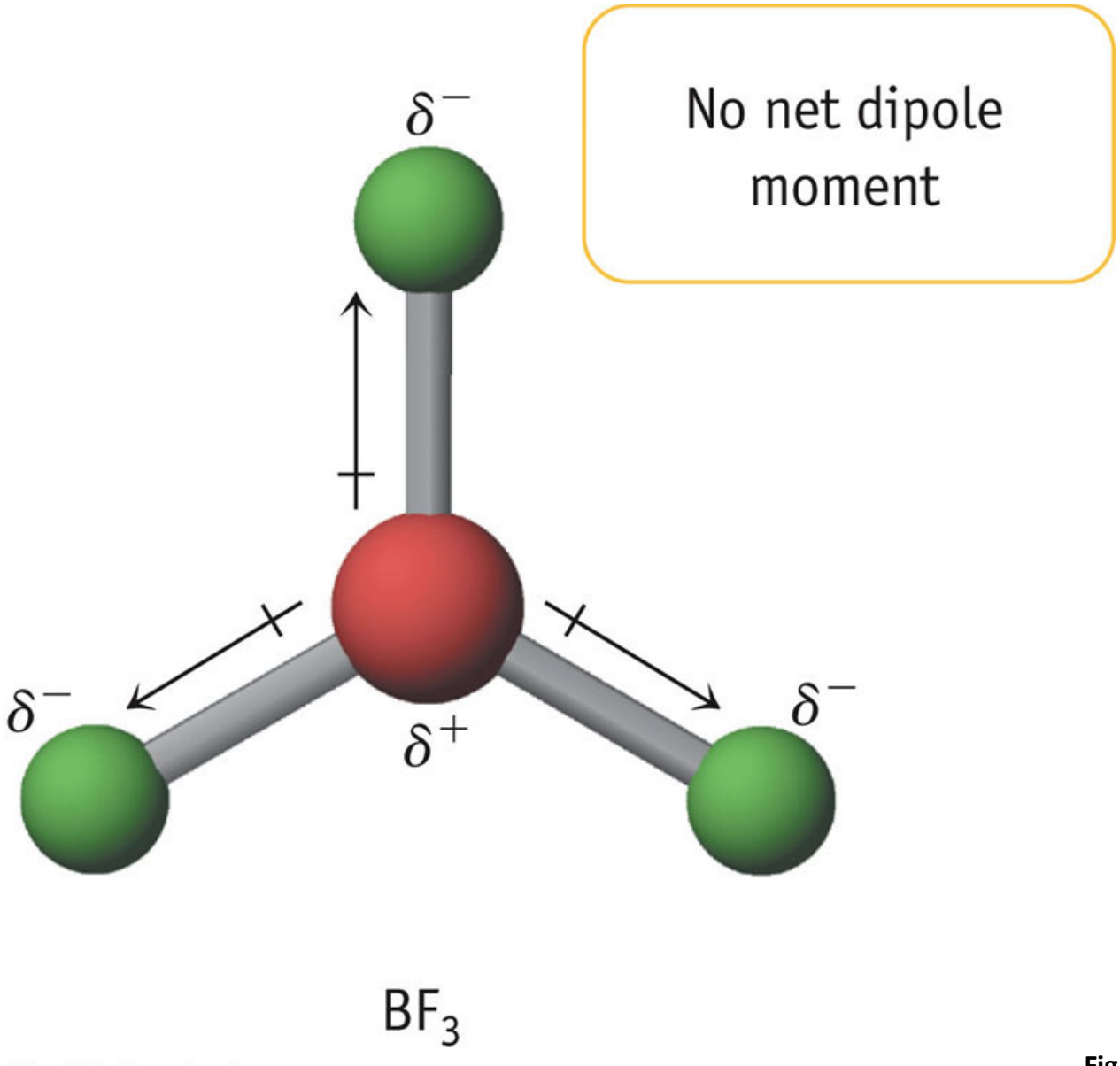


Fig. 8-14, p. 383

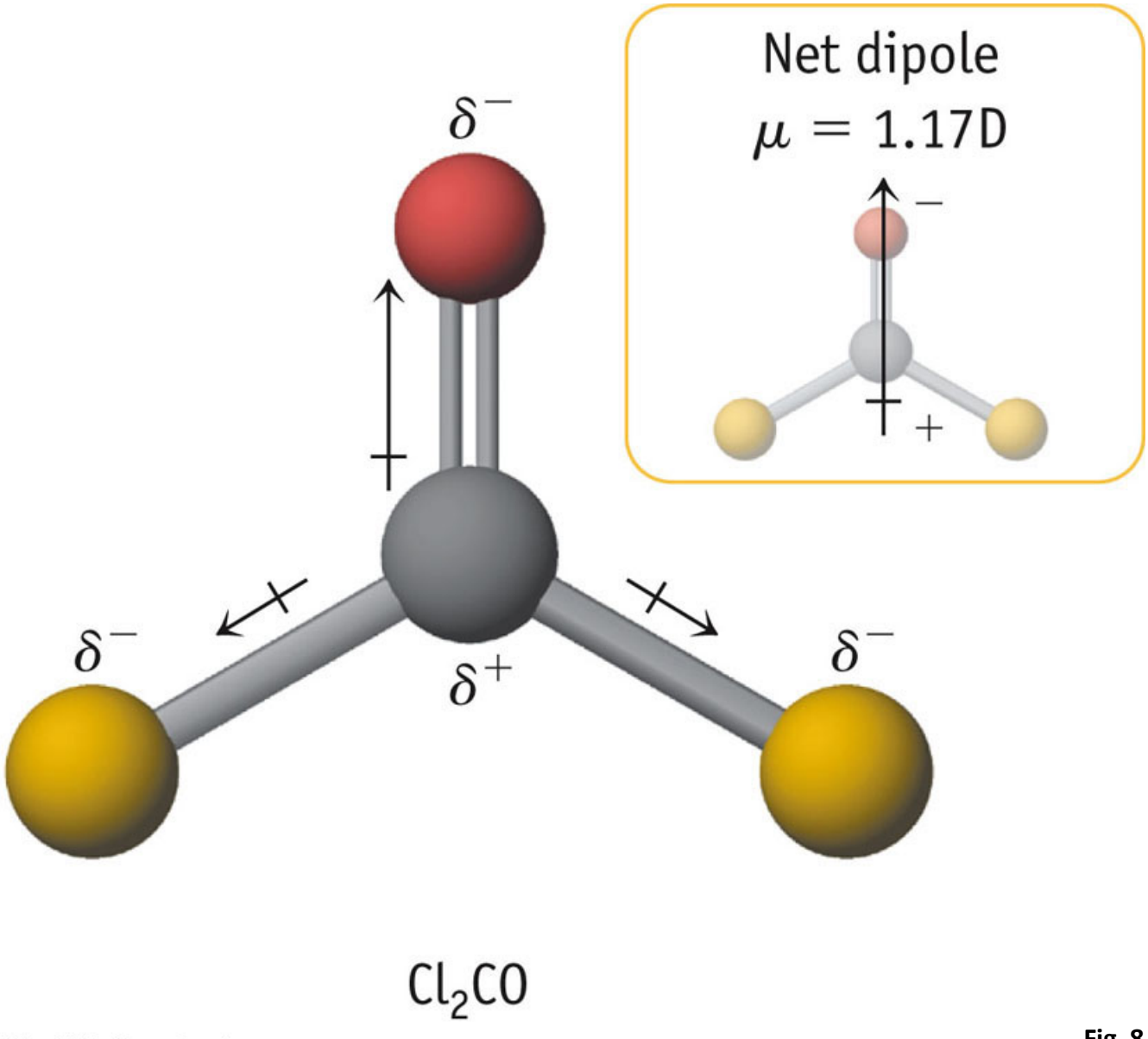


Fig. 8-14, p. 383

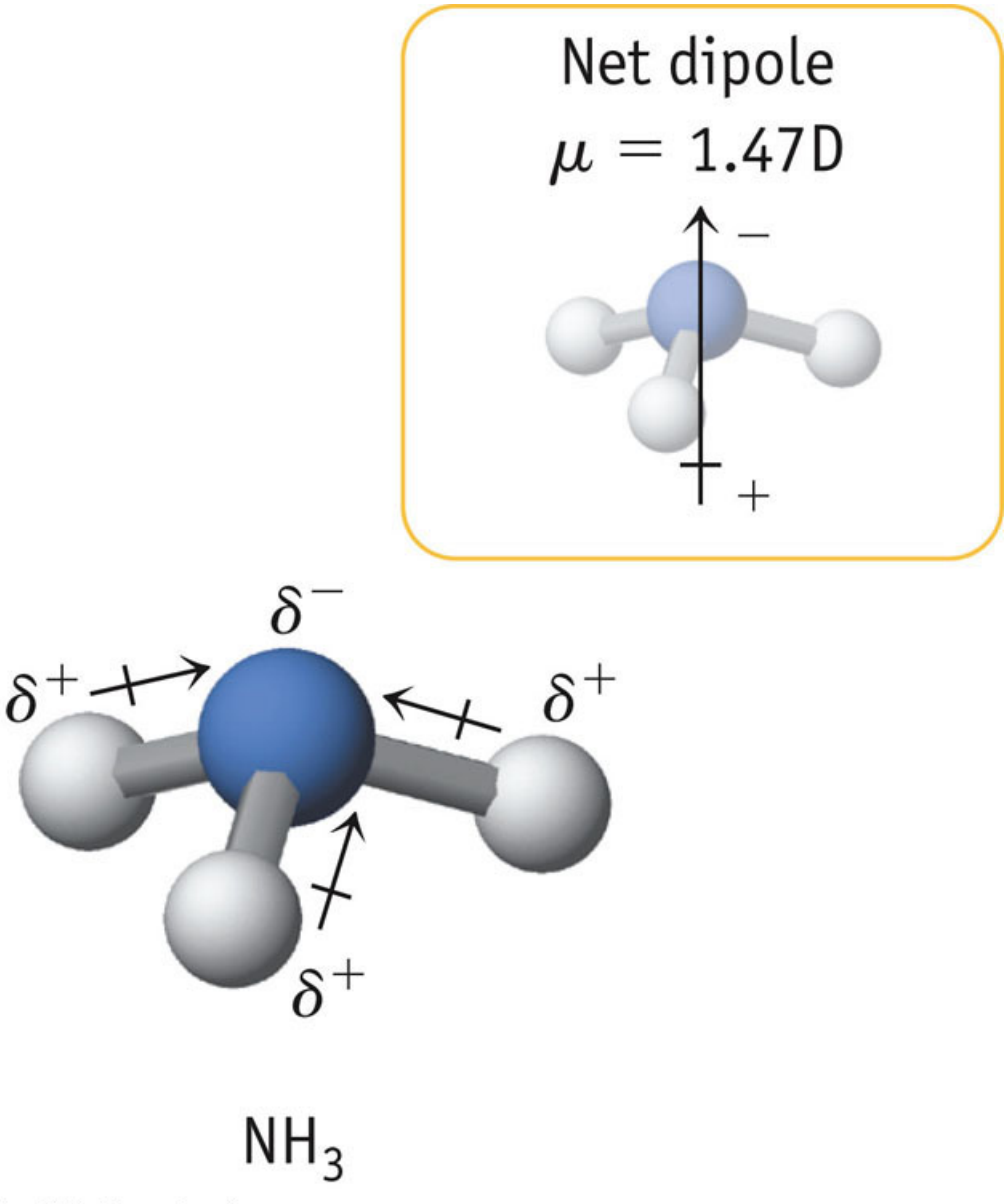
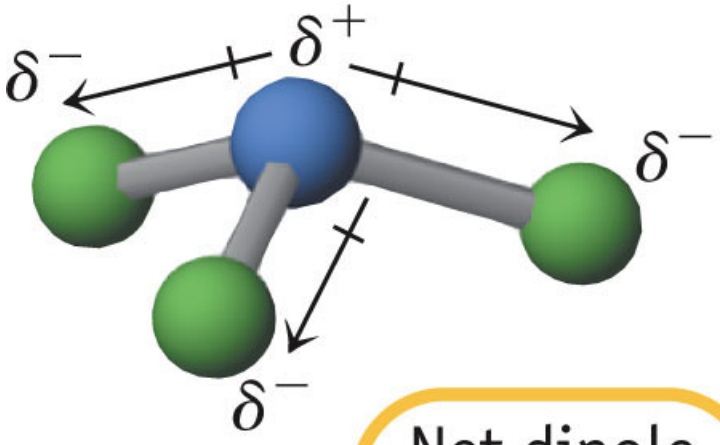
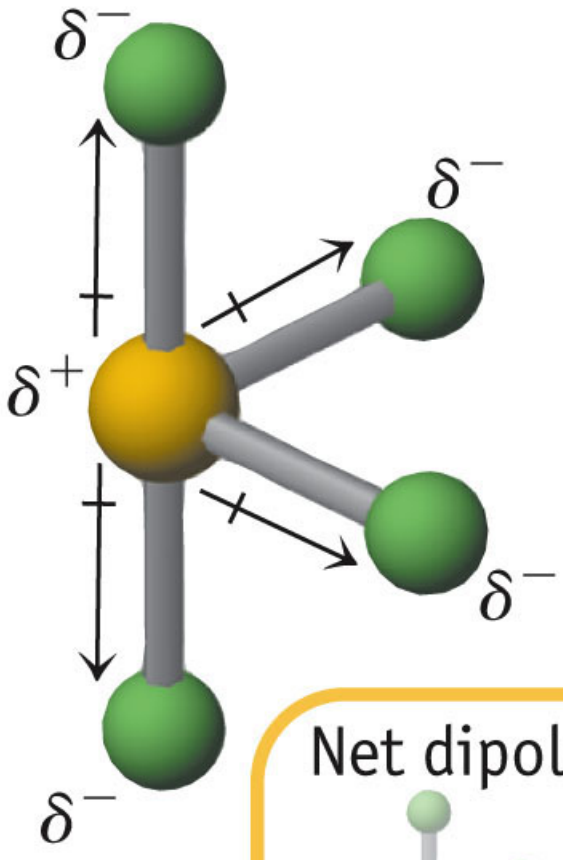
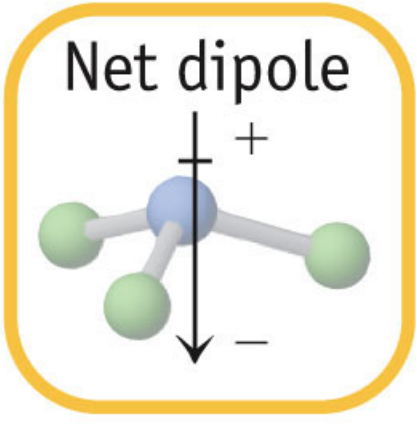


Fig. 8-14, p. 383

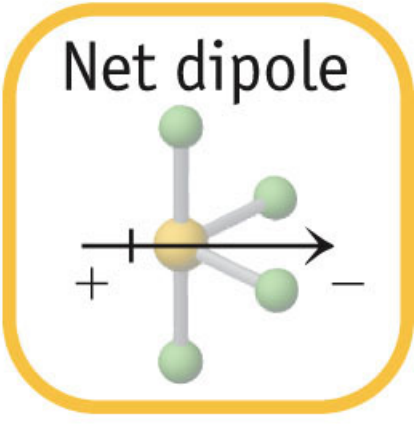




NF_3



SF_4



Example 8-12, p. 384

TABLE 8.8 Some Average Single- and Multiple-Bond Lengths in Picometers (pm)*

Single Bond Lengths

	Group 1A	4A	5A	6A	7A	4A	5A	6A	7A	7A	7A
	H	C	N	O	F	Si	P	S	Cl	Br	I
H	74	110	98	94	92	145	138	132	127	142	161
C		154	147	143	141	194	187	181	176	191	210
N			140	136	134	187	180	174	169	184	203
O				132	130	183	176	170	165	180	199
F					128	181	174	168	163	178	197
Si						234	227	221	216	231	250
P							220	214	209	224	243
S								208	203	218	237
Cl									200	213	232
Br										228	247
I											266

Multiple Bond Lengths

C═C	134	C≡C	121
C═N	127	C≡N	115
C═O	122	C≡O	113
N═O	115	N≡O	108

*1 pm = 10^{-12} m.

Table 8-8, p. 387

TABLE 8.9 Some Average Bond Dissociation Enthalpies (kJ/mol)*

Single Bonds

	H	C	N	O	F	Si	P	S	Cl	Br	I
H	436	413	391	463	565	328	322	347	432	366	299
C		346	305	358	485	—	—	272	339	285	213
N			163	201	283	—	—	—	192	—	—
O				146	—	452	335	—	218	201	201
F					155	565	490	284	253	249	278
Si						222	—	293	381	310	234
P							201	—	326	—	184
S								226	255	—	—
Cl									242	216	208
Br										193	175
I											151

Multiple Bonds

N═N	418	C═C	610
N≡N	945	C≡C	835
C═N	615	C═O	745
C≡N	887	C≡O	1046
O═O (in O_2)	498		

*Sources of dissociation enthalpies: I. Klotz and R. M. Rosenberg: *Chemical Thermodynamics*, 4th Ed., p. 55, New York, John Wiley, 1994; and J. E. Huheey, E. A. Keiter, and R. L. Keiter: *Inorganic Chemistry* 4th Ed., Table E. 1, New York, Harper-Collins, 1993. See also Lange's *Handbook of Chemistry*, J. A. Dean (ed.), McGraw-Hill Inc., New York.

Table 8-9, p. 389