Bonding

Ionic bond is what holds ionic compounds together

(1) start w/ neutral atoms $Na_{(g)}$ and $Cl_{(g)}$

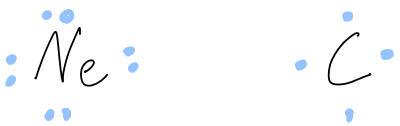
(2) transfer of valence e^- $Na^+_{(g)}$ and $Cl^-_{(g)}$

(3) oppositely charged ions attract $NaCl$

ionic bond = electrostatic interaction btw ions

Covalent bonding involves sharing of valence e^-

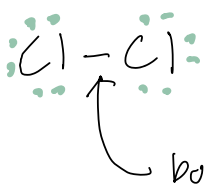
Use Lewis structure to represent
- valence electrons described with a dot
- most atoms want "octet" of valence electrons (noble gas config)



if the 2 atoms get close enough they can share e^-

each Cl has 6 of "own" e^- and 2 shared e^- → octet

Cl_2:



Making "good" Lewis structures

(1) Determine total # of valence e^-

(2) Determine atom arrangement/connectivity
often central vs peripheral atoms
H, F, Cl, Br, I tend to only form 1 bond
O, S, Se tend to form 2 bonds
N, P tend to form 3 bonds
C, Si tend to form 4 bonds

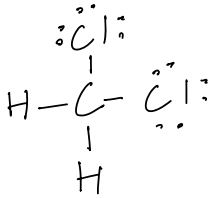
(3) Connect all atoms using covalent bonds (= $2e^-$)

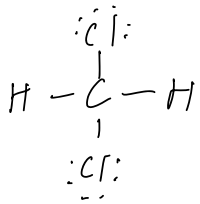
(4) Distribute remaining e^- as lone pairs to get atoms to octet * H wants $2e^-$ not $8e^-$

(5) Check - e^-, octet

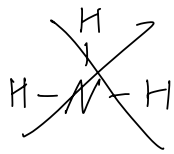
(6) If some atoms "unhappy" consider additional bond

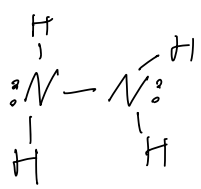
CH_2Cl_2
4 2x1 2x7 = 20 e^-



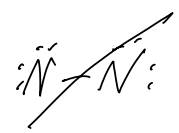


N_2H_4
2x5 4x1 = 14 e^-

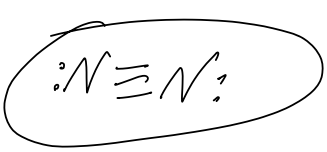




N_2 - 10 e^-

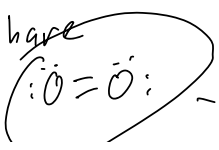

share more


share more


triple bond

O_2 - 12 e^-


share


- double bond

CO - 10 e^- :C≡O:

Formal charge (of an atom)

bisect all bonds, "give" 1 e⁻ to each atom

add to #e⁻ in lone pairs for each atom

compare to # of valence e⁻ lone atom would have

:N≡N:

↓ ↓

∅ ∅

(2+3) - 5

:C≡O:

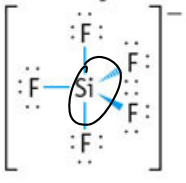
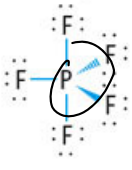
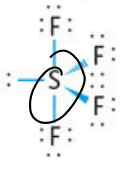
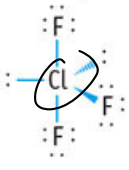
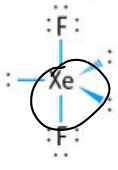
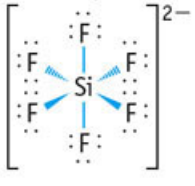
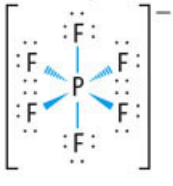
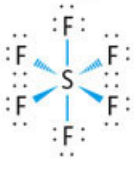
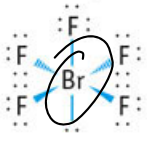
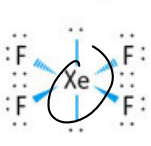
↓ ↓

-1 +1

Non-zero Formal charge → reactive molecule

farther down periodic table

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

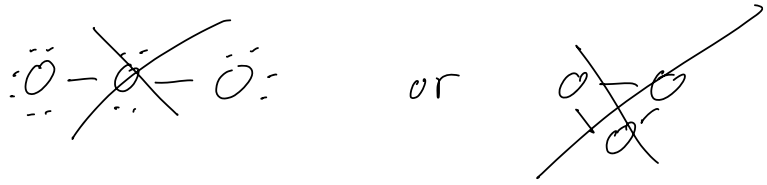
Group 4A	Group 5A	Group 6A	Group 7A	Group 8
SiF_5^- 	PF_5 	SF_4 	ClF_3 	XeF_2 
SiF_6^{2-} 	PF_6^- 	SF_6 	BrF_5 	XeF_4 

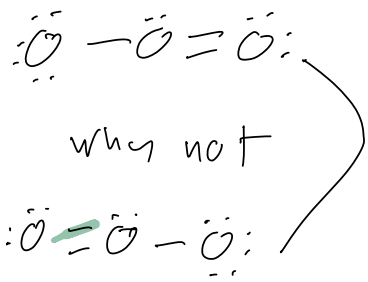
H only needs 2e

B only needs 6e

O_3 - ozone

18e⁻



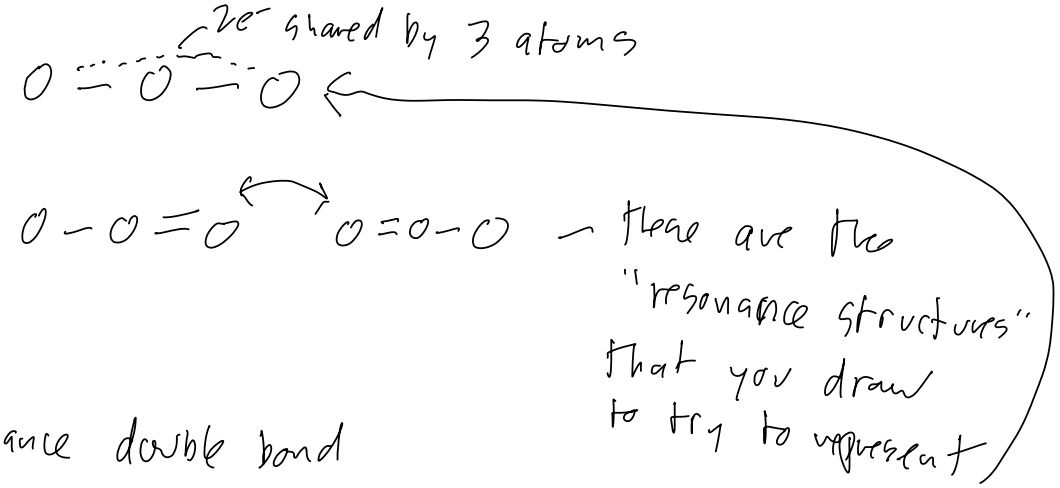


why not

These are actually the same

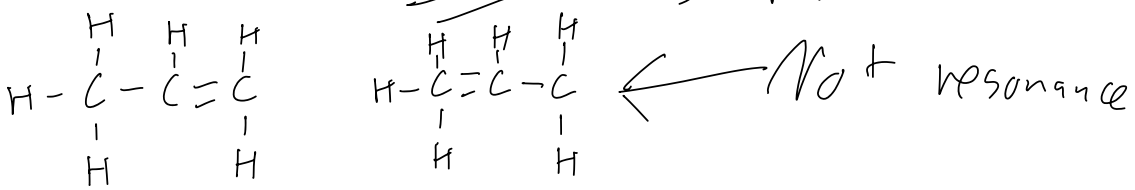
* You'd expect an asymmetric molecule
* Expts show a symmetric molecule

"Resonance" - 2e⁻ in "double bond" shared by all 3 O atoms in molecule



These are the "resonance structures" that you draw to try to represent

For resonance double bond needs to be able to be put in different places without moving atom



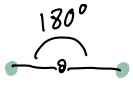
Not resonance

Valence Shell Electron Pair Repulsion (VSEPR)
to describe 3D-ness of molecules
* electron pairs want to avoid each other

Electron pair geometry

2 e⁻ pairs

Linear

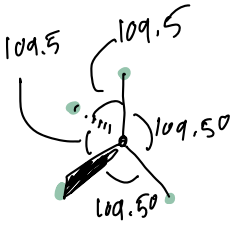


3 e⁻ pairs

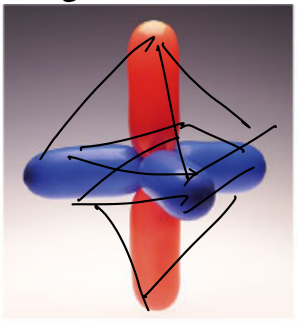
Trigonal planar

4 e⁻ pair

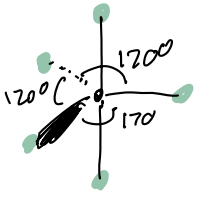
Tetrahedral



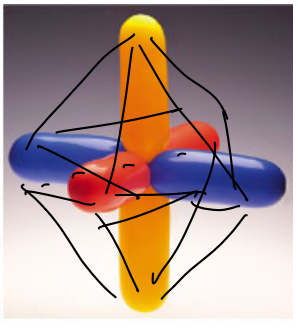
5 e⁻ pairs



Trigonal bipyramidal



6 e⁻ pair



Octahedral

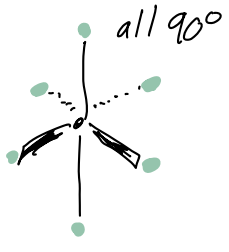


Fig. 8-4, p. 368

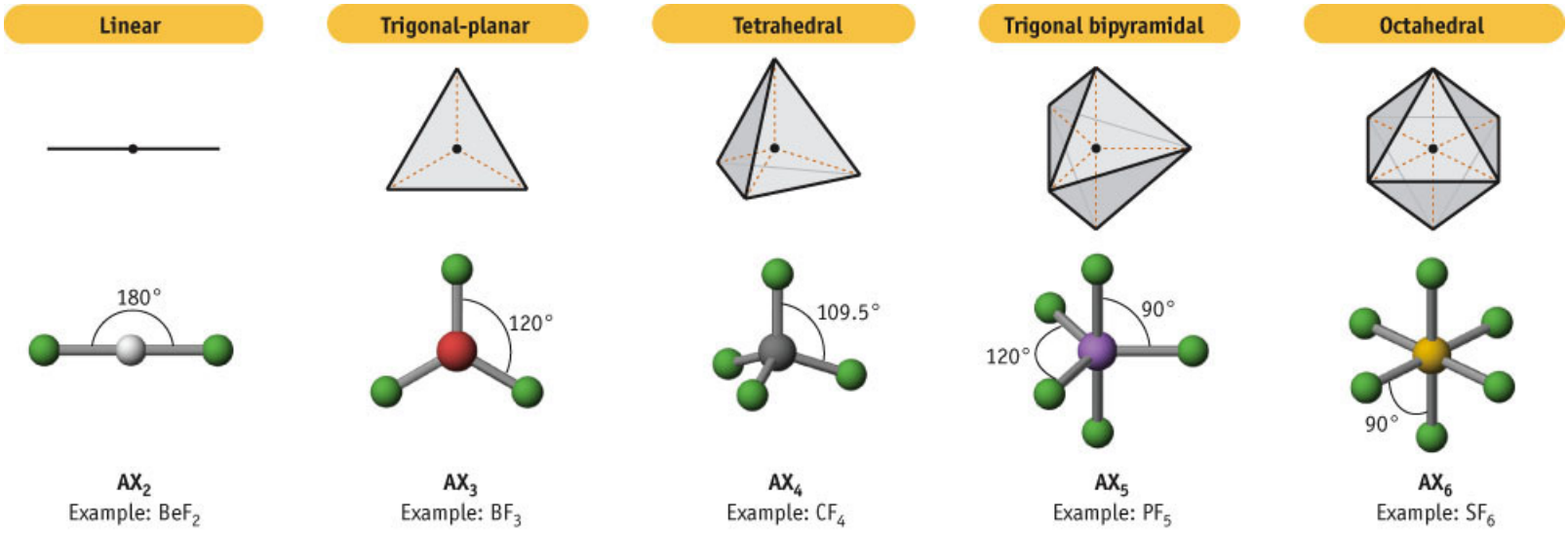


Fig. 8-5, p. 369

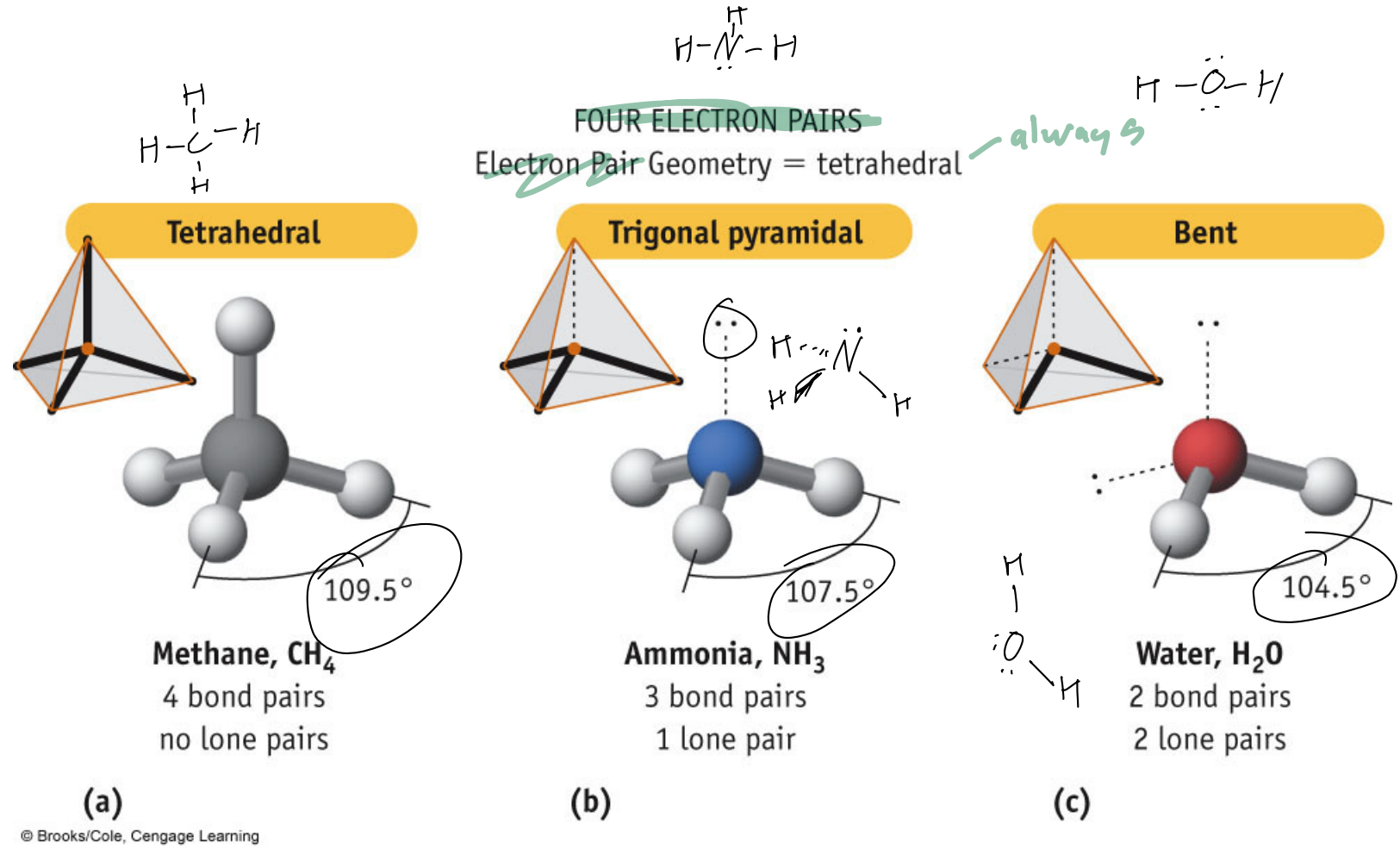
consider each atom with >1 thing connected to it

determine # of electron pairs around it → electron pair geometry

figure out bonds vs lone pairs → molecular geometry

electron pair geometry not necc same as molec. geometry

molecular geometry – where atoms are



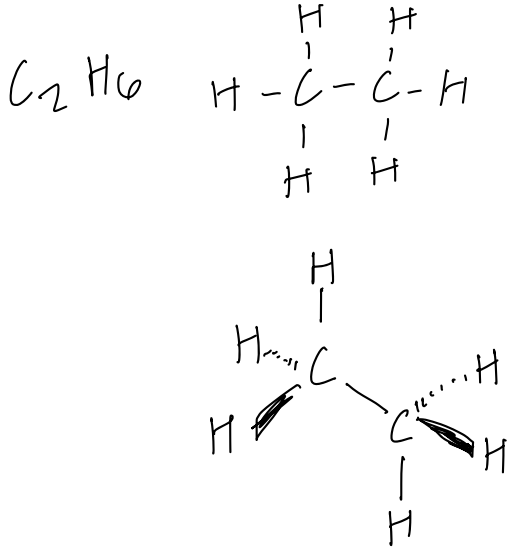


Fig. 8-6, p. 371

5 e⁻ pairs → trig bipyramid

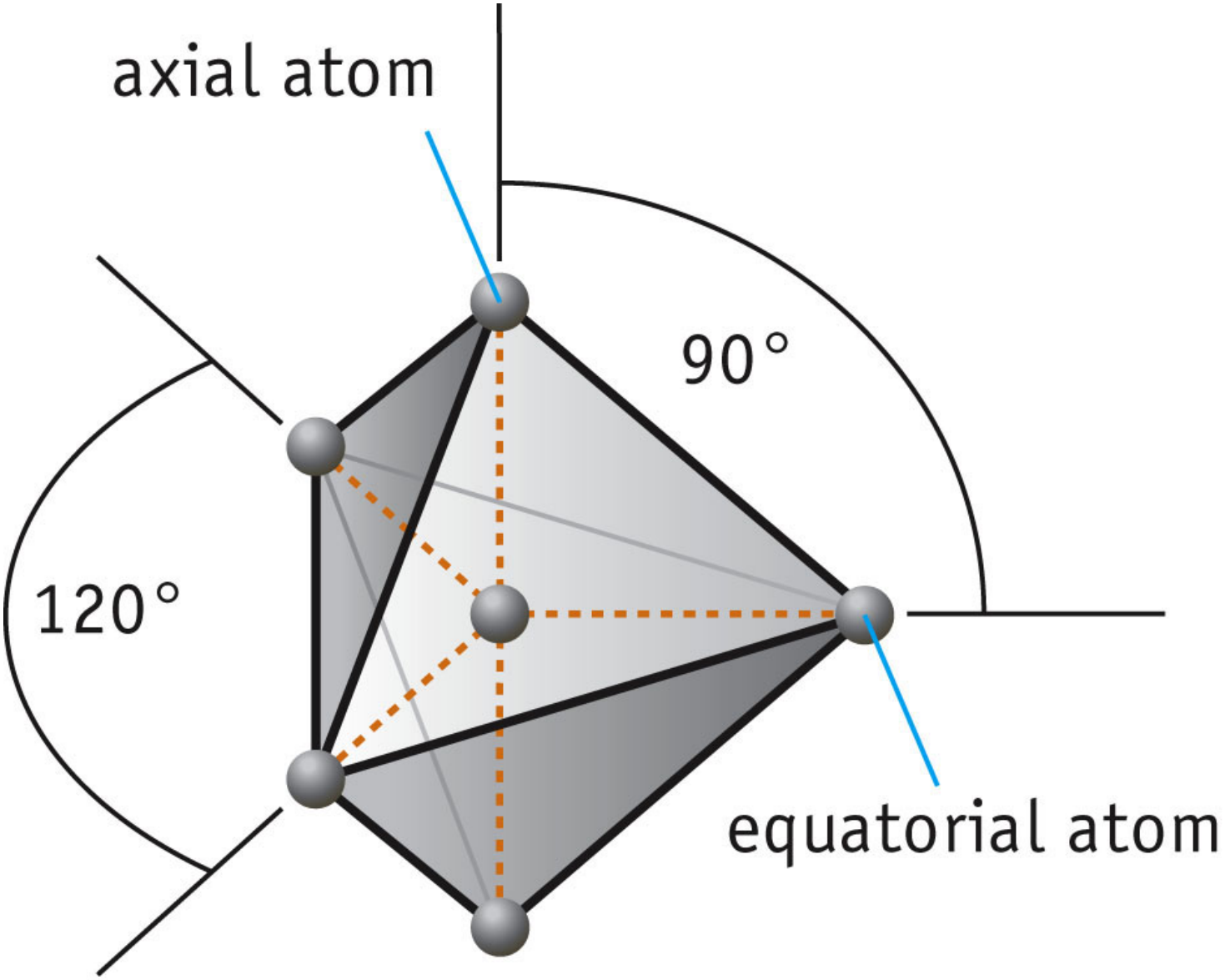
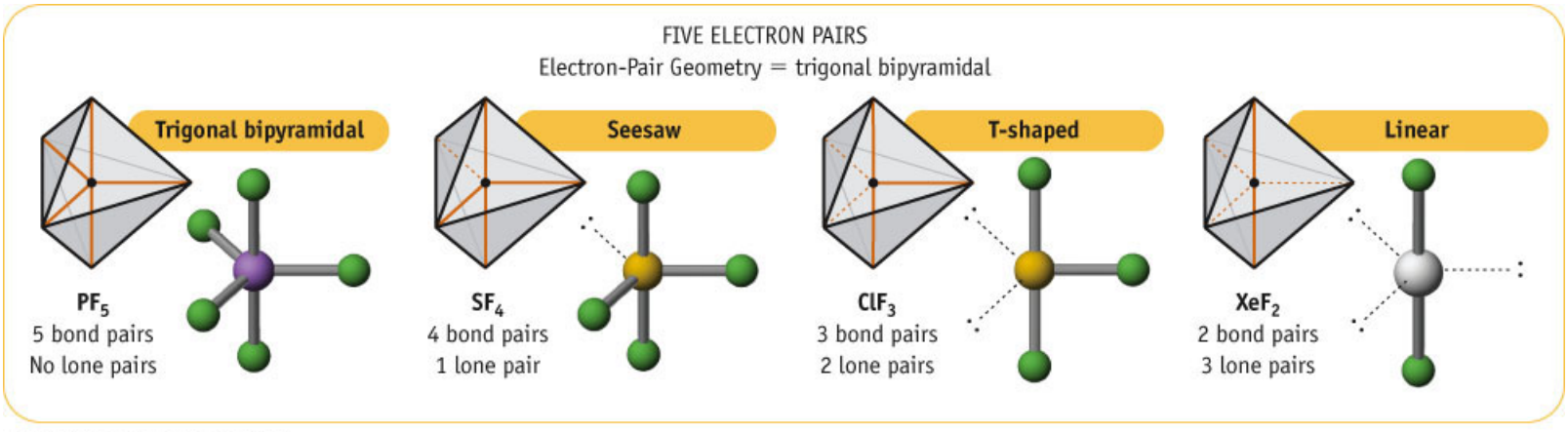


Fig. 8-7, p. 372



Swap on lone pairs to equitorial position

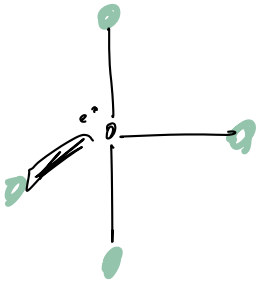


Fig. 8-8, p. 372

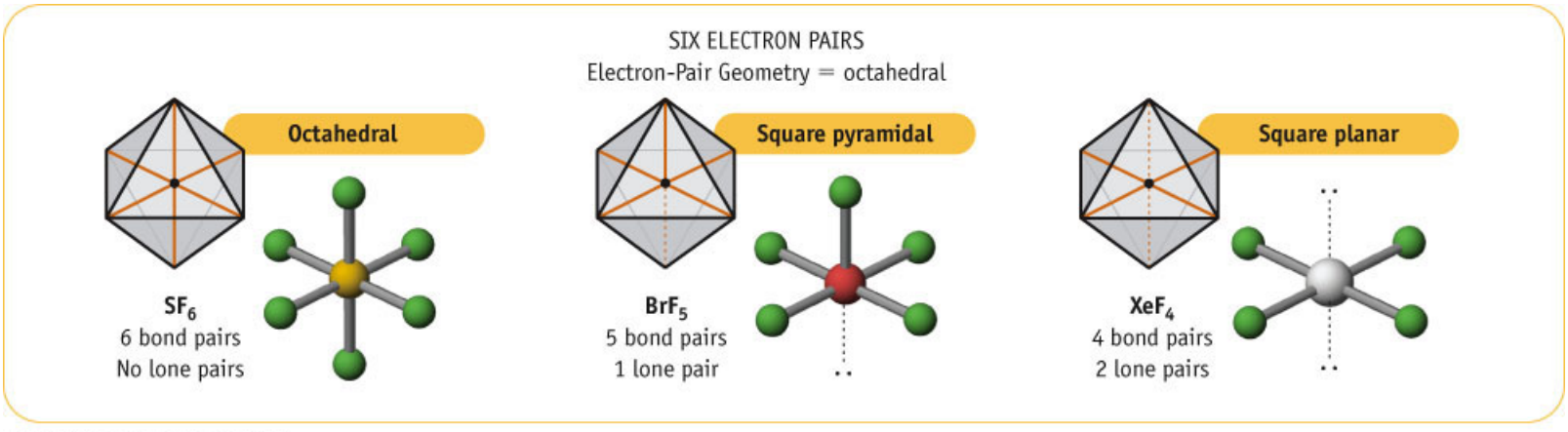


Fig. 8-8, p. 372

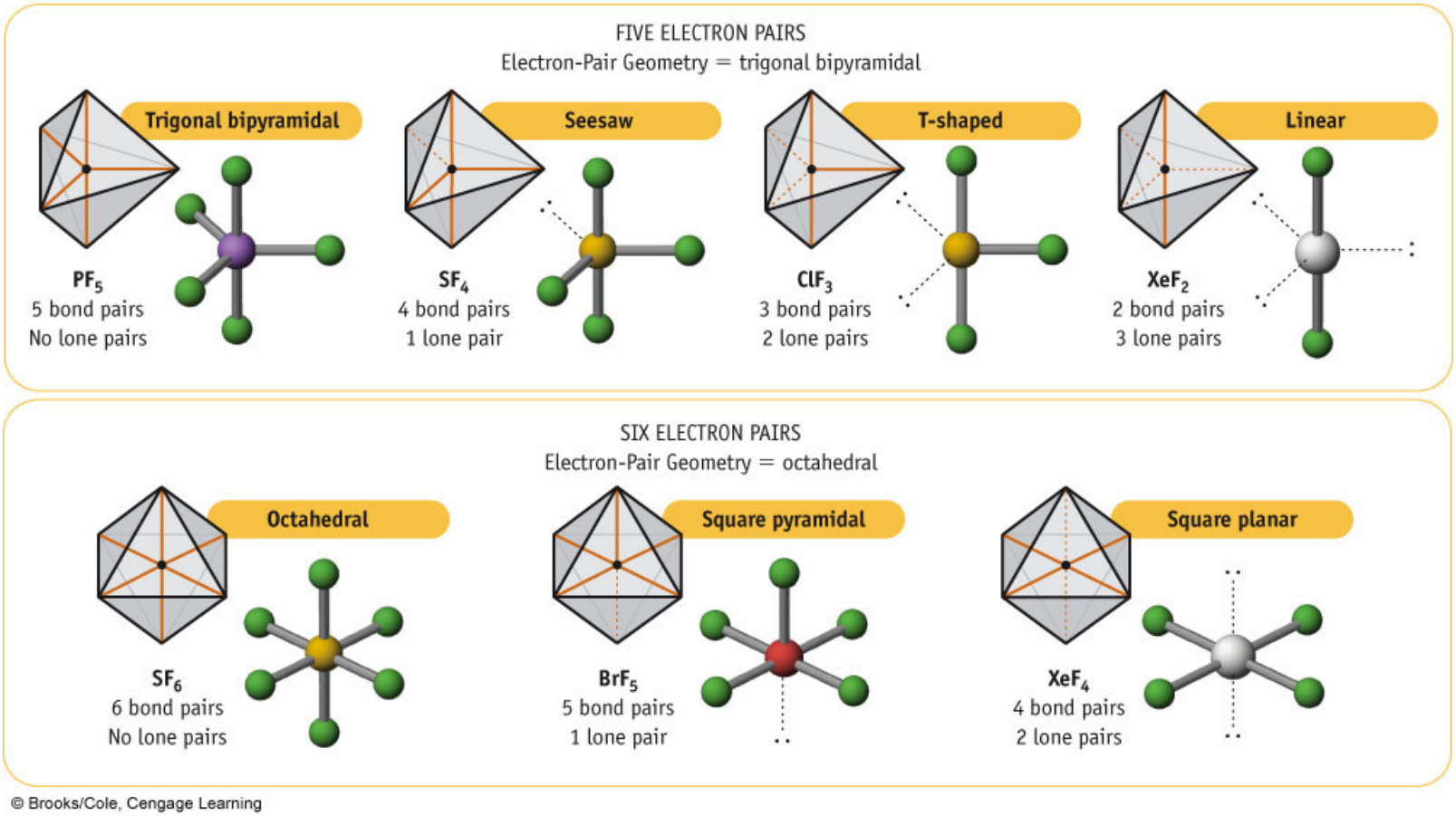
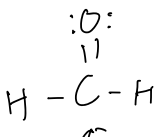


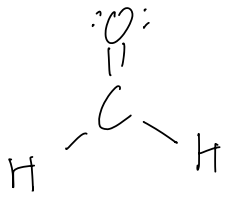
Fig. 8-8, p. 372

Treat double/triple bond as one "superpair" of electrons for VSEPR purposes

CH_2O



treat as if 3 e⁻ pairs not 4



"Ideal" covalent bond has equitable sharing of electrons

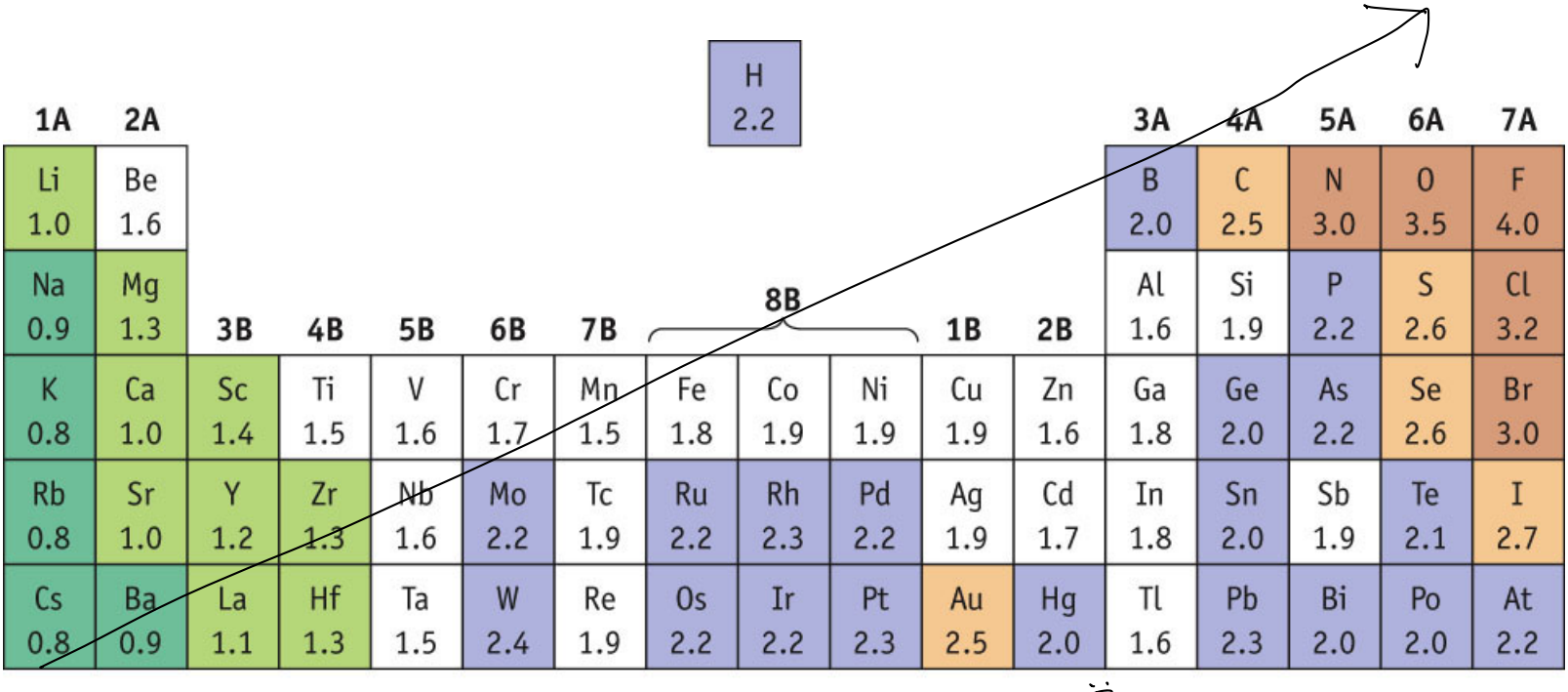
often there is unequal sharing of those 2 e⁻ resulting in one end of bond w/ more electron density and other end w/ less e⁻ density

"Polar bond"

Electronegativity - how "selfish" atom is about sharing electrons

Δ electroneg determines polarity of a bond

If Δ is ≥ 0.5 polar bond



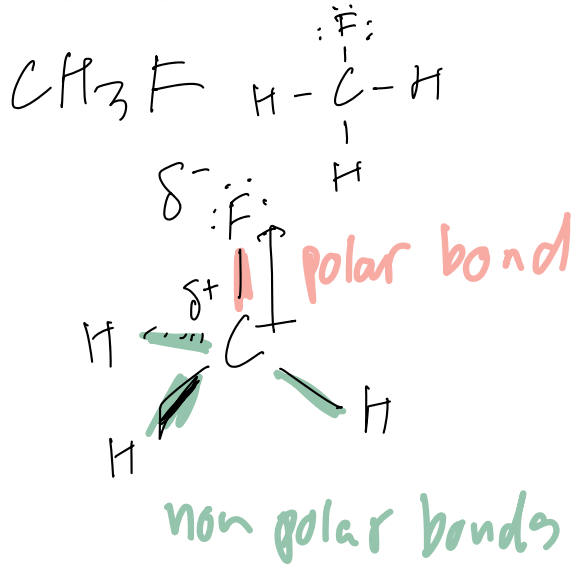


Fig. 8-11, p. 376

Can also have polar molecules

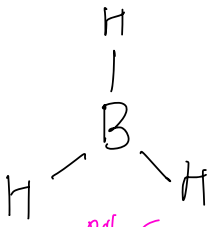
B = 2.0
H = 2.2
F = 4.0

BH_3

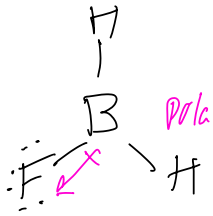
BH_2F

BHF_2

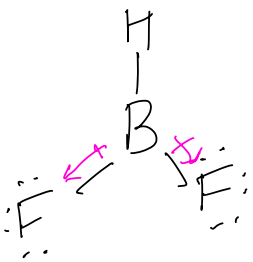
BF_3



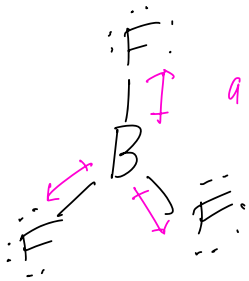
non polar molec



polar molecule

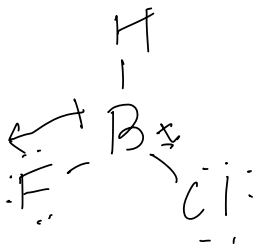


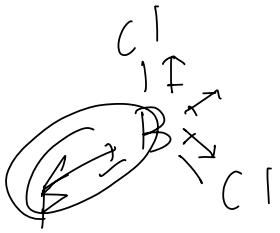
polar



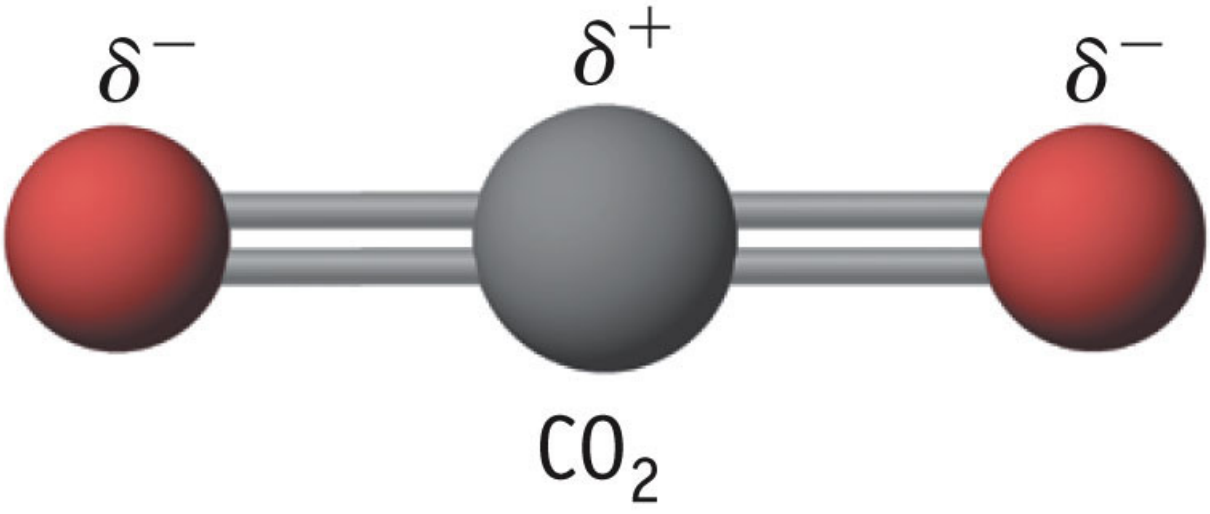
all vectors cancel

non polar molecule



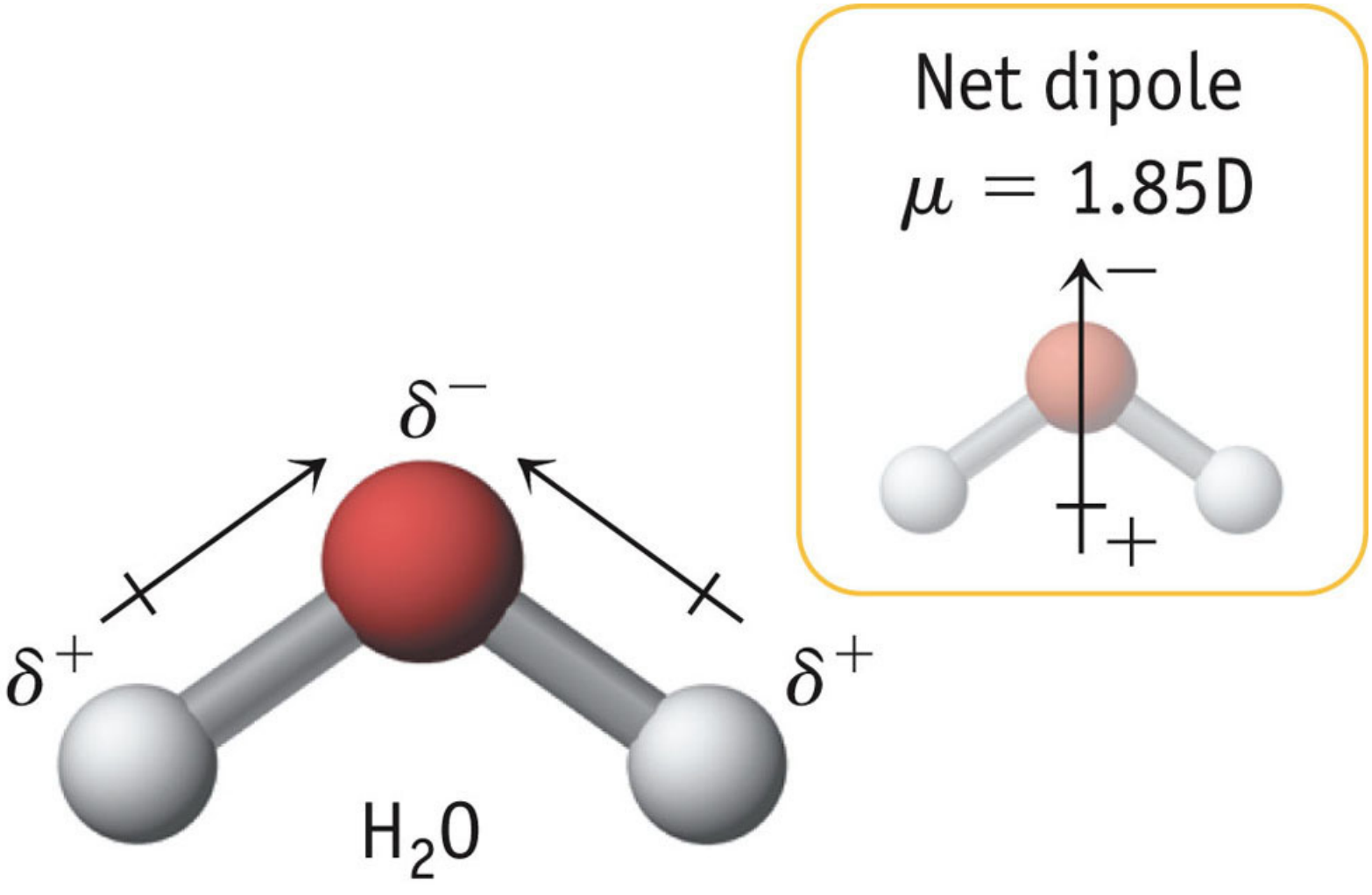


No net dipole moment



(a)

Fig. 8-13, p. 381



(b)

Fig. 8-13, p. 381

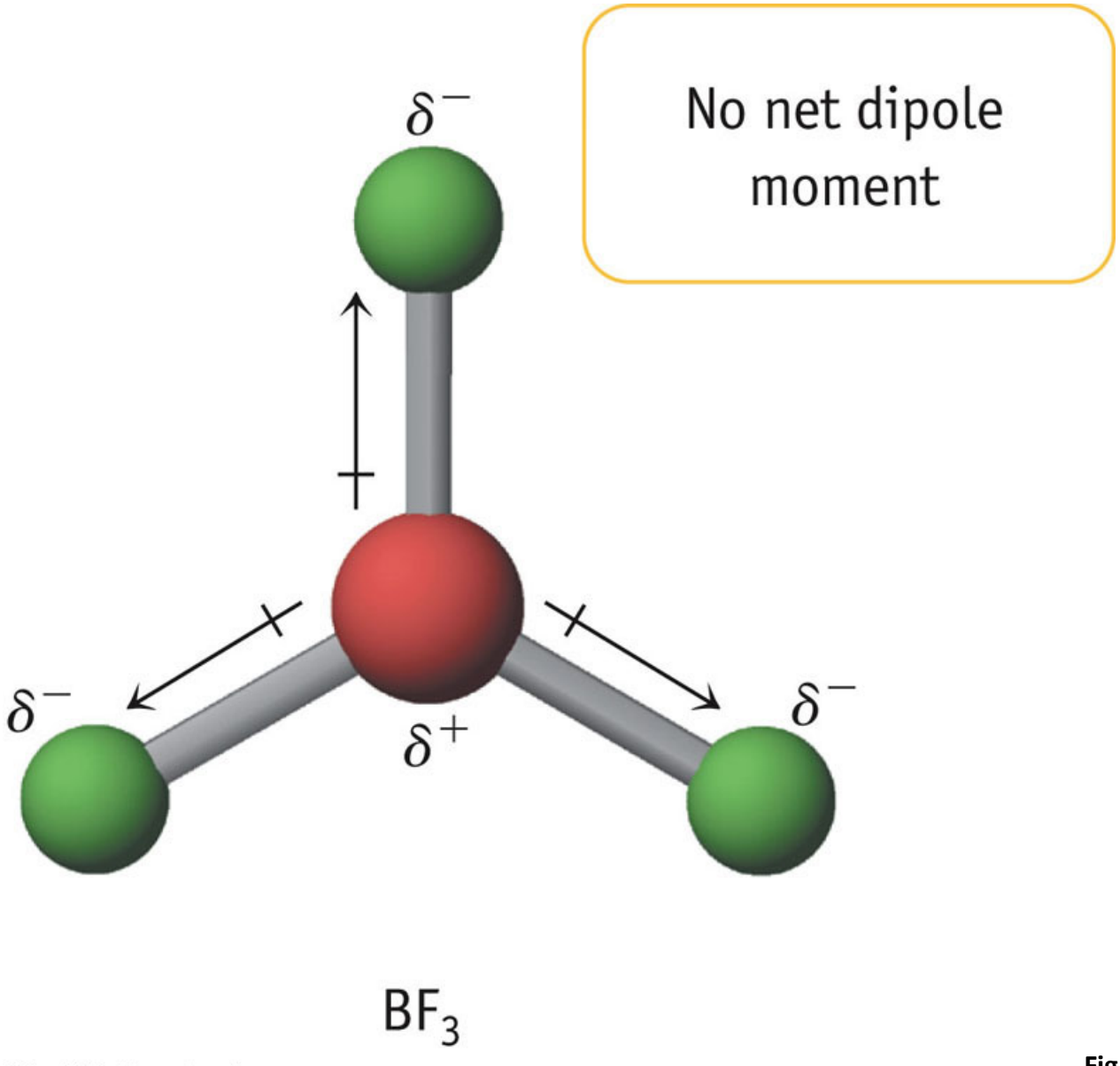


Fig. 8-14, p. 383

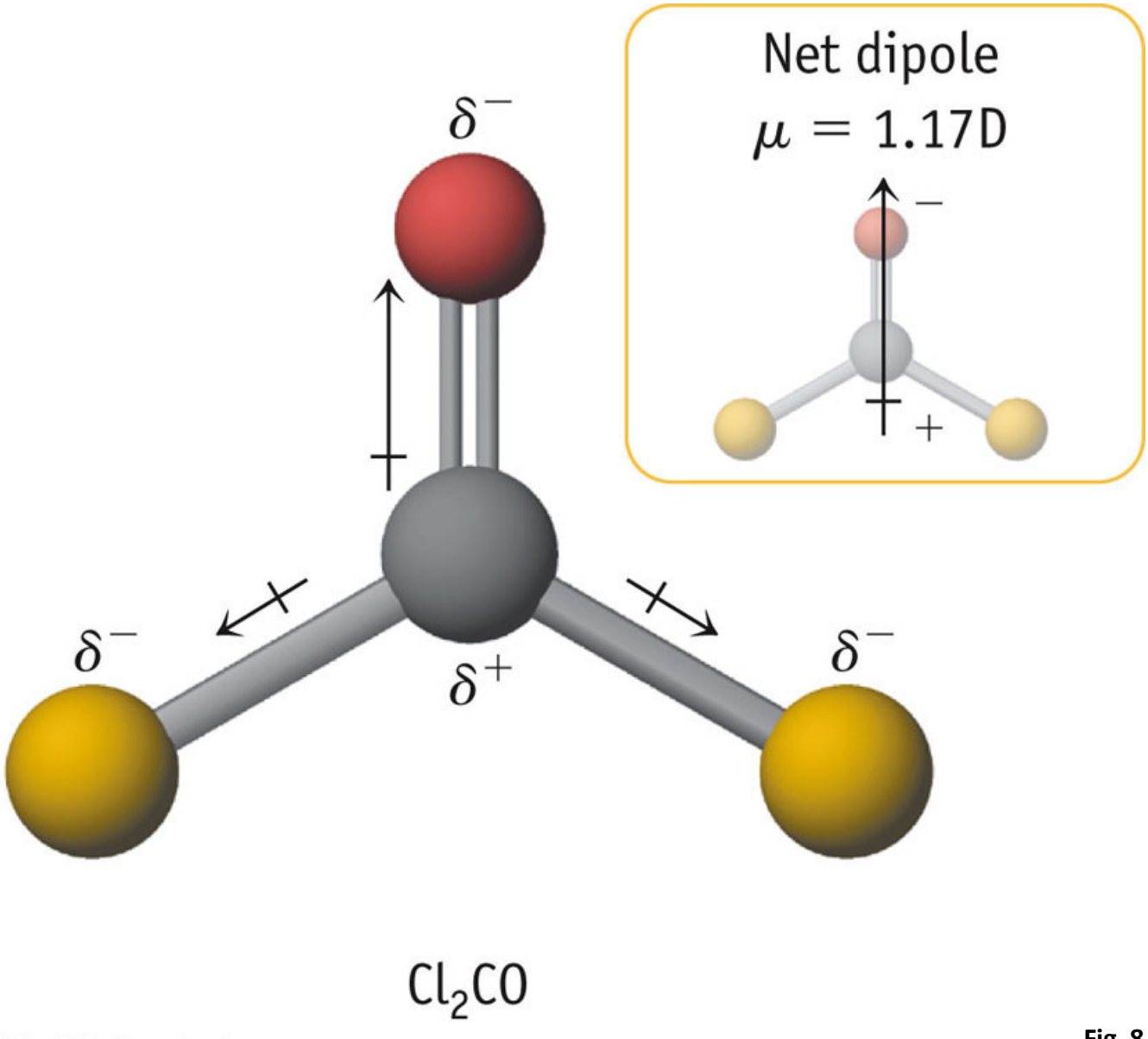


Fig. 8-14, p. 383

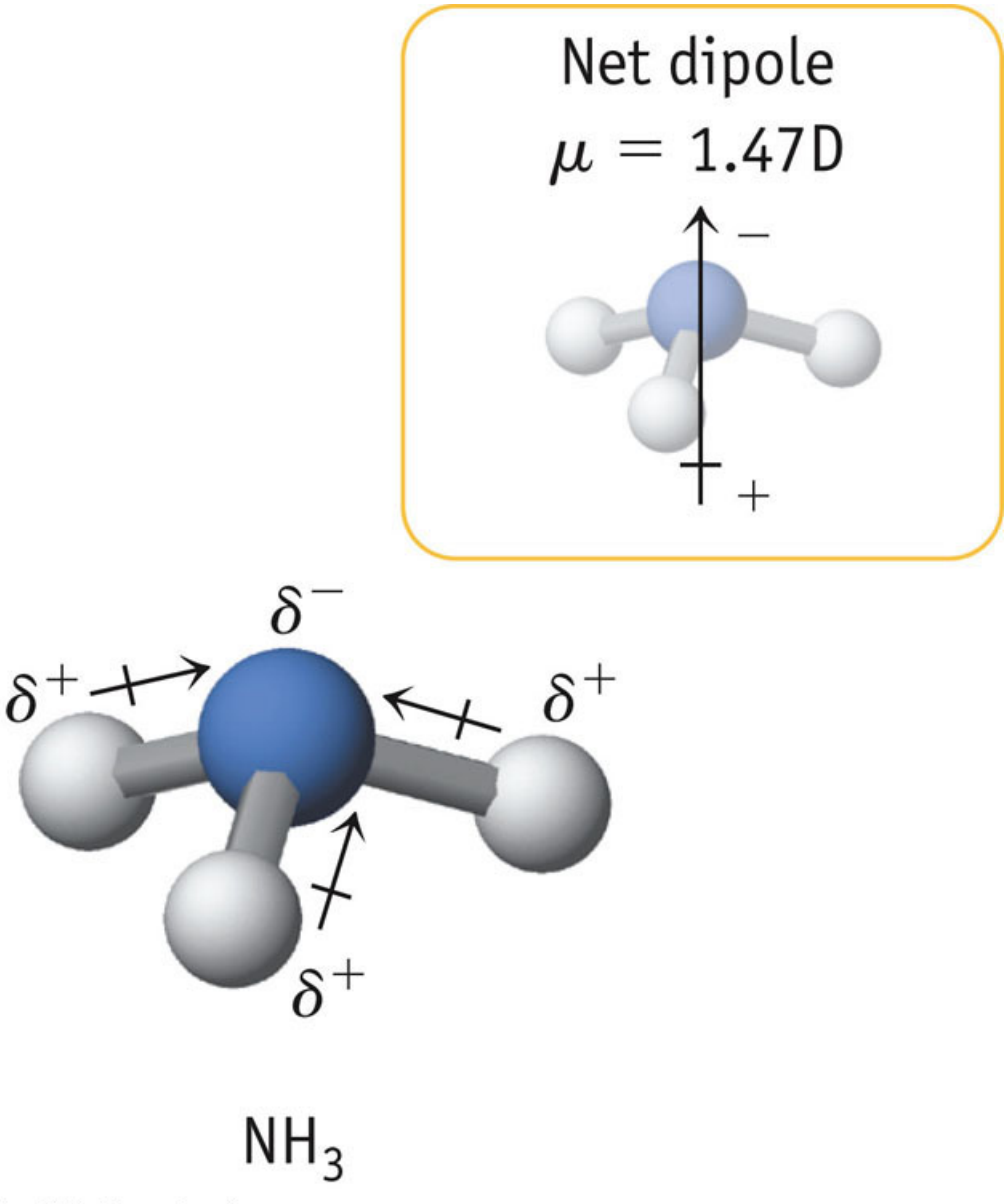
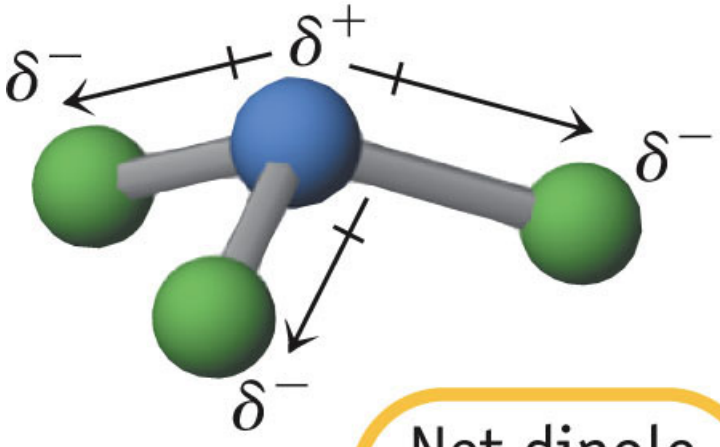
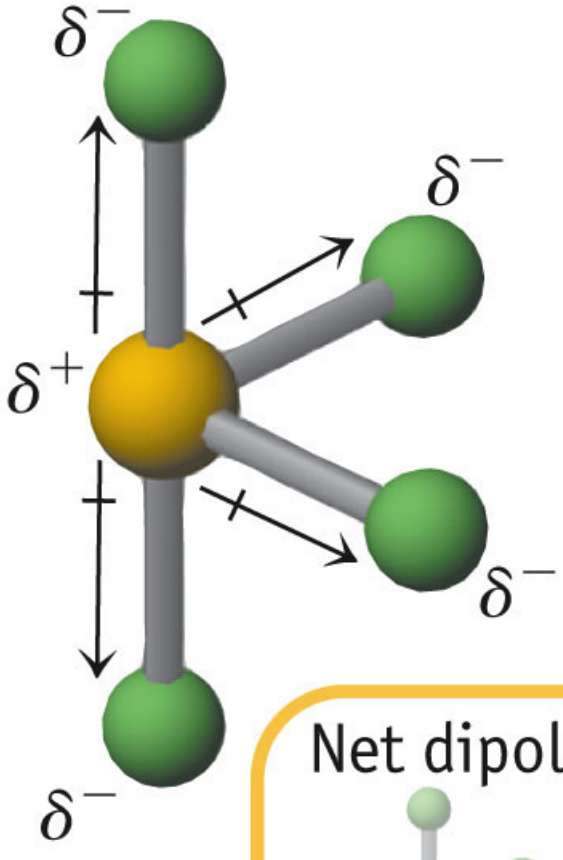
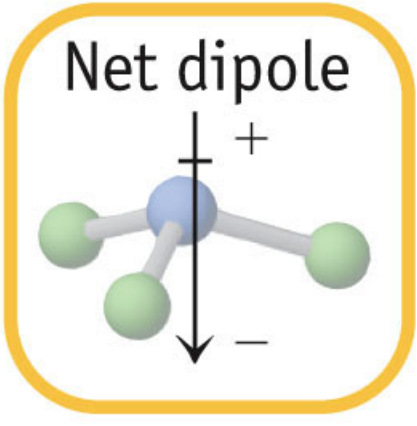


Fig. 8-14, p. 383

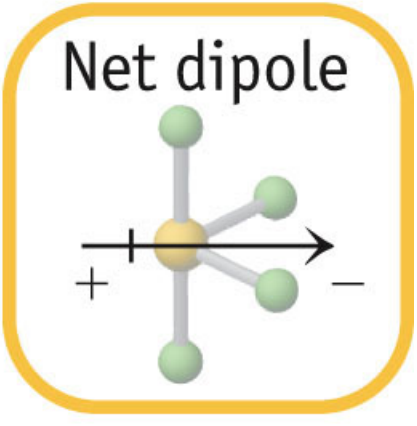




NF_3



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