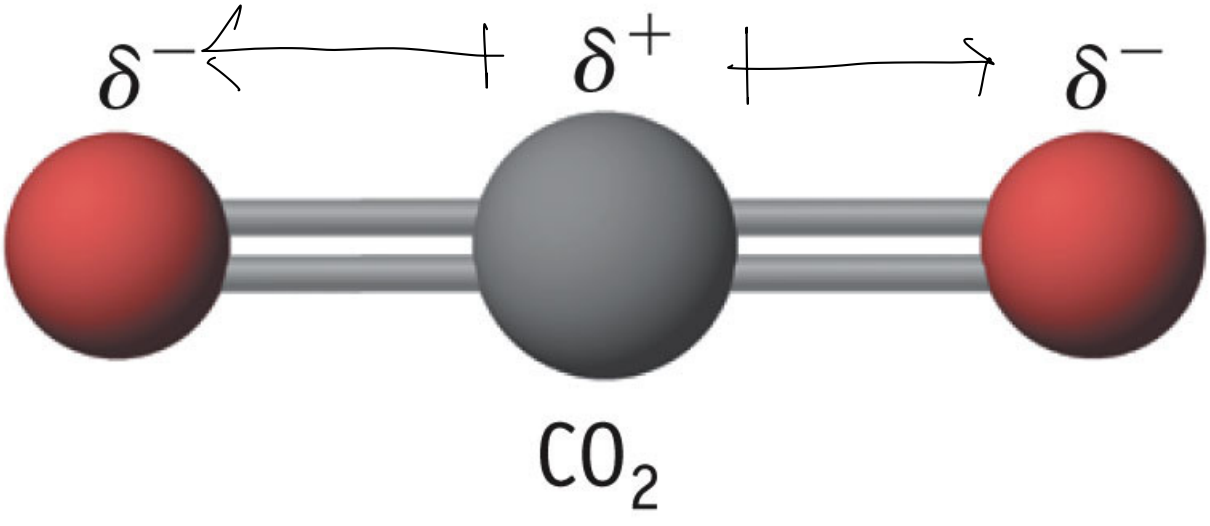
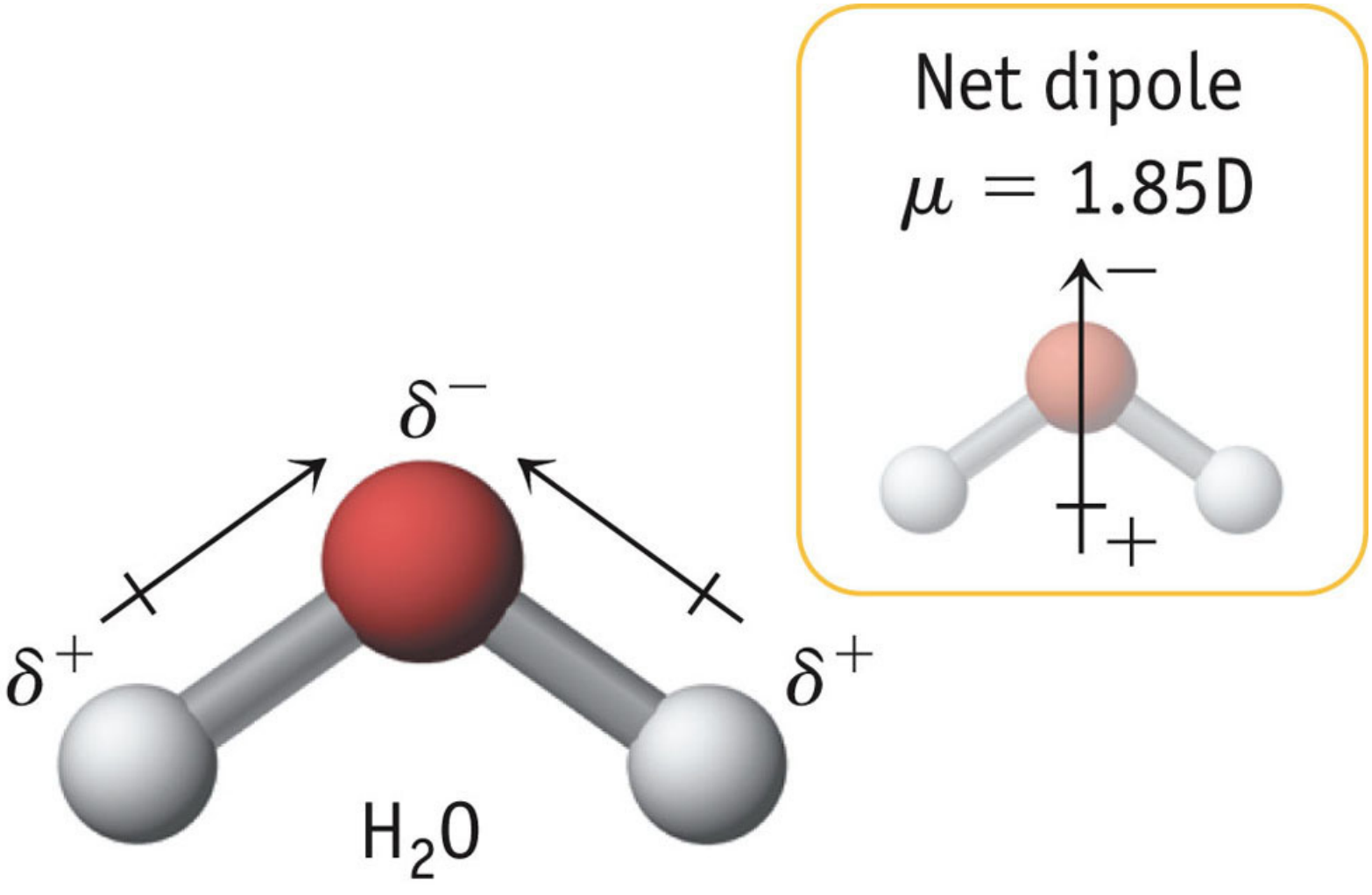
No net dipole moment



(a)

Fig. 8-13, p. 381



(b)

Fig. 8-13, p. 381

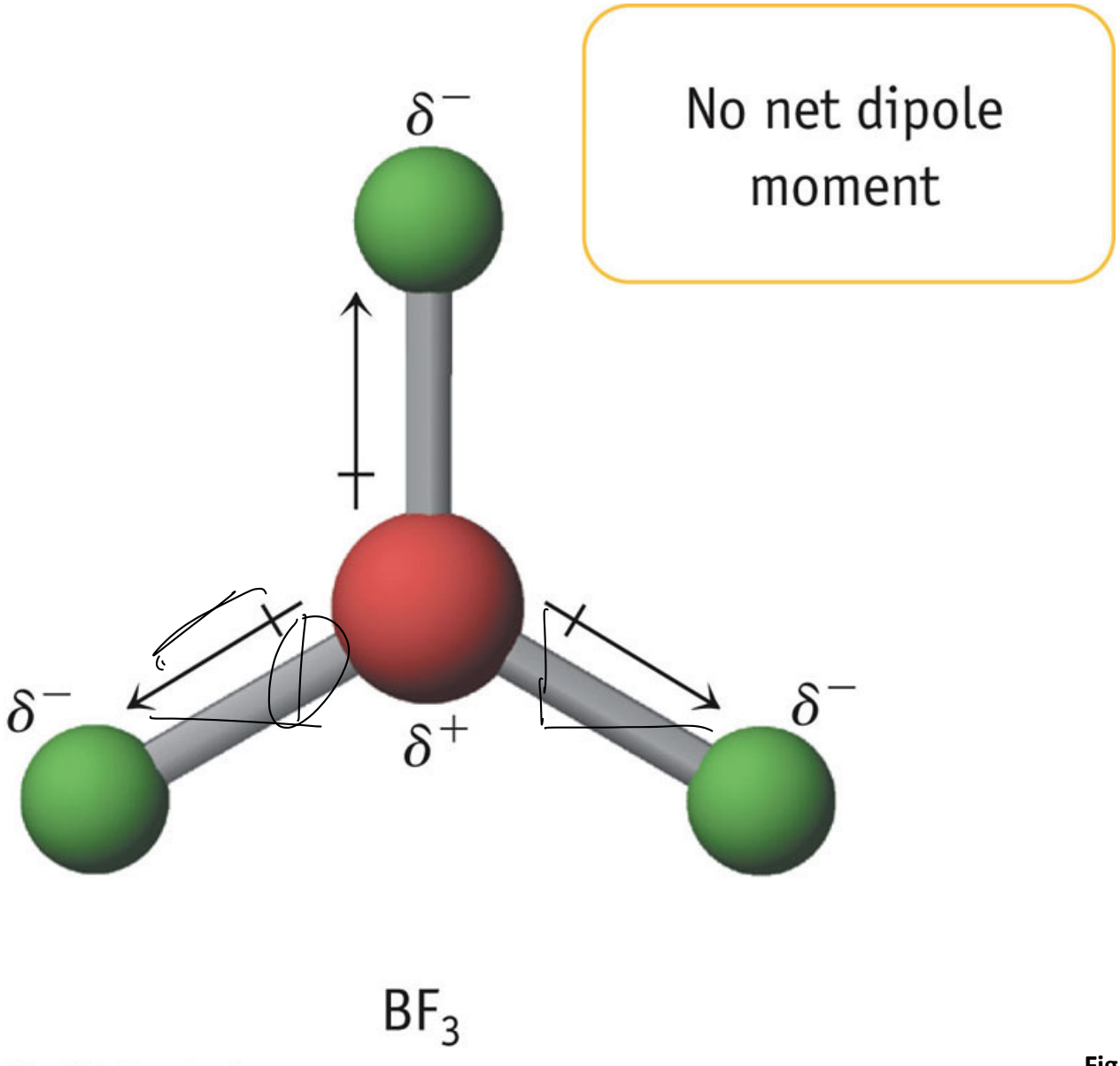


Fig. 8-14, p. 383

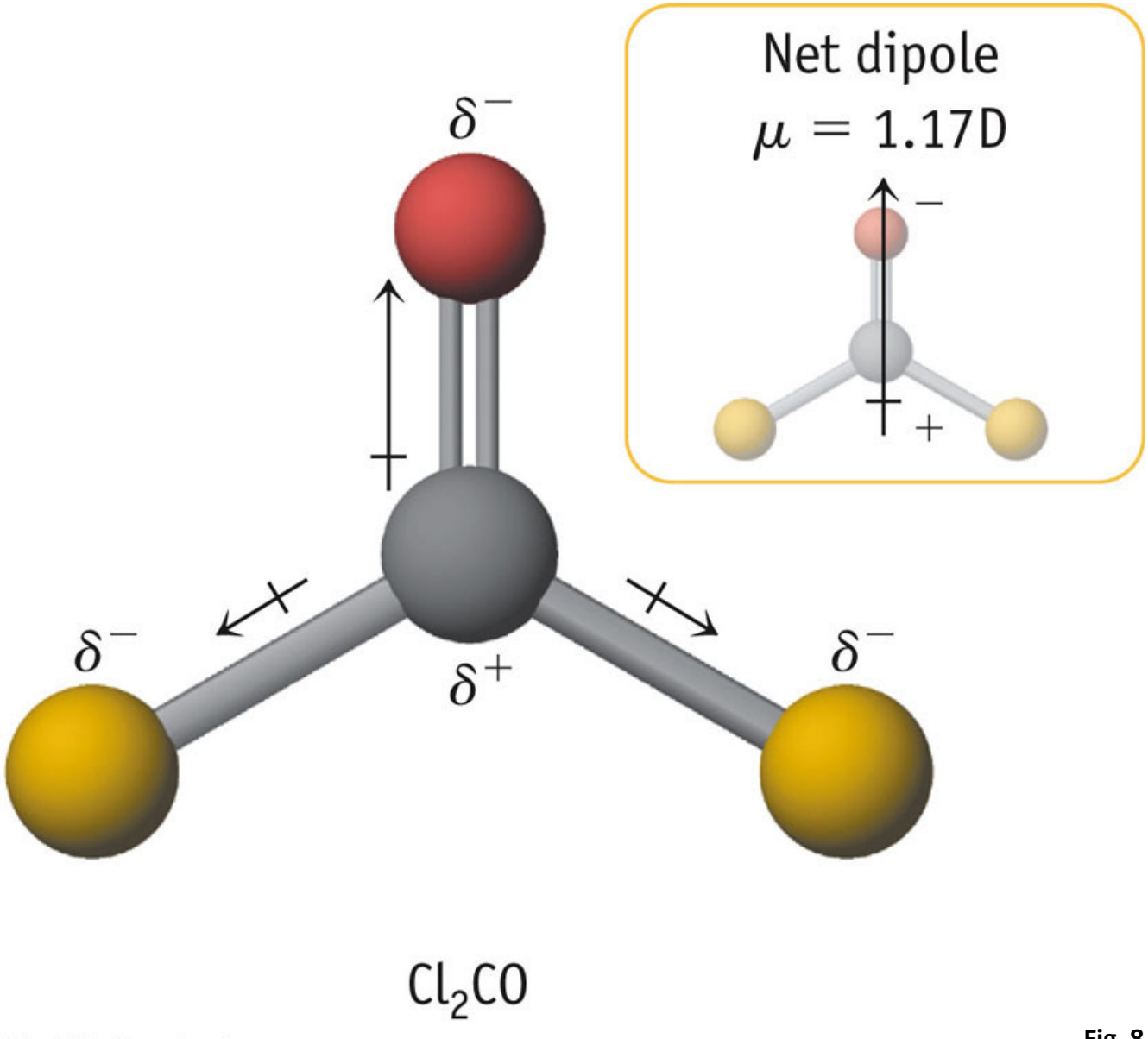


Fig. 8-14, p. 383

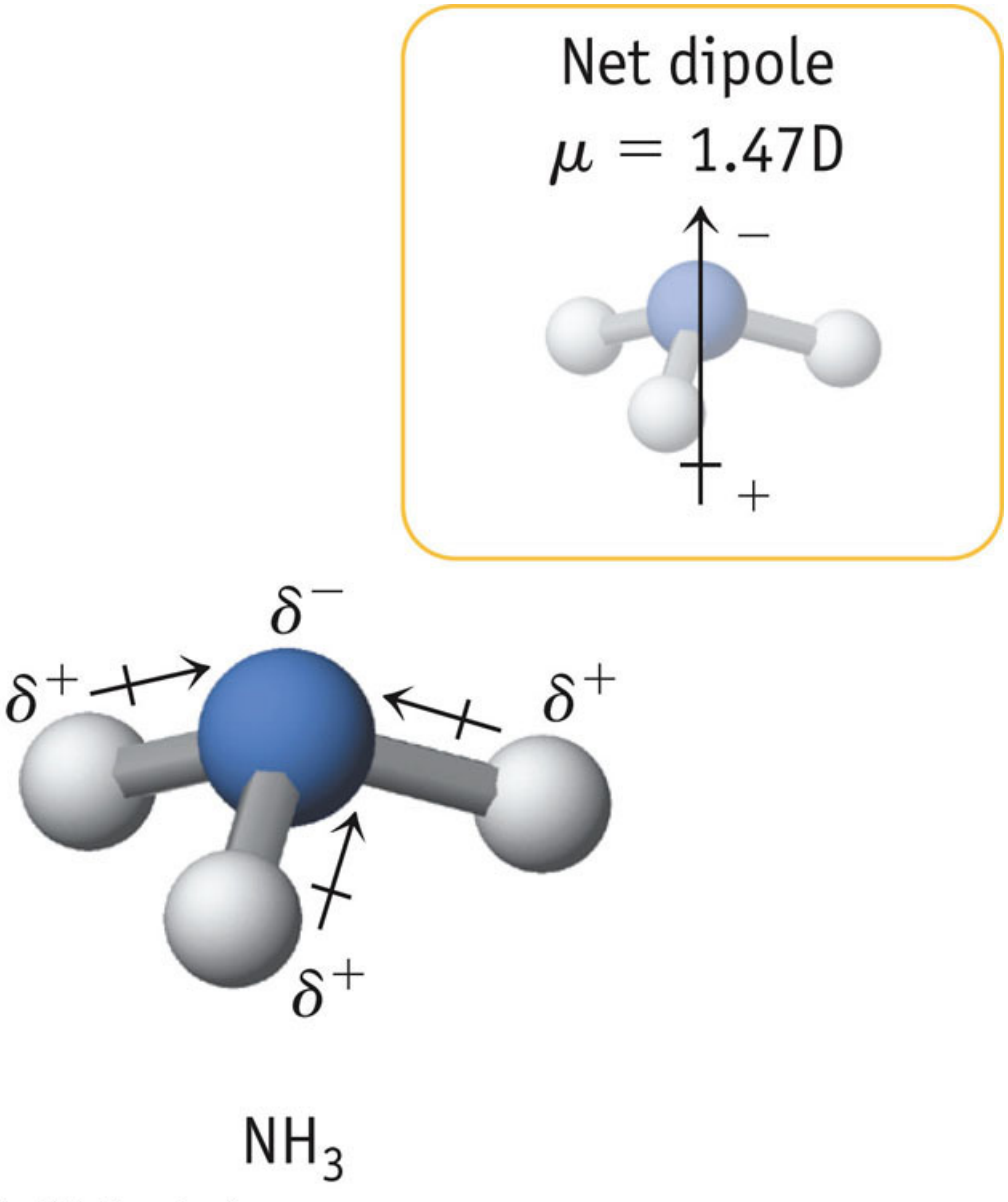
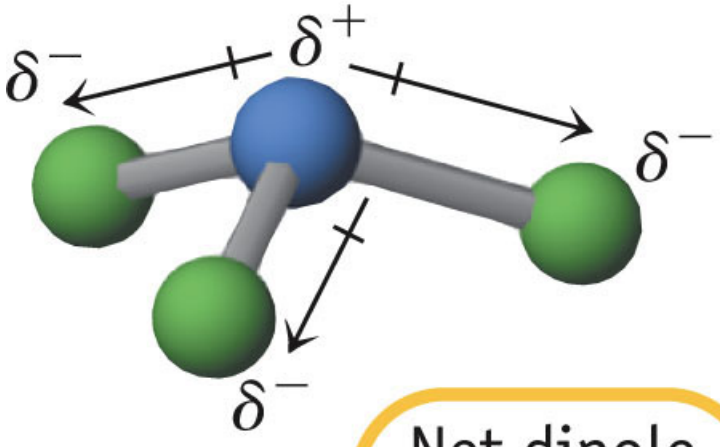
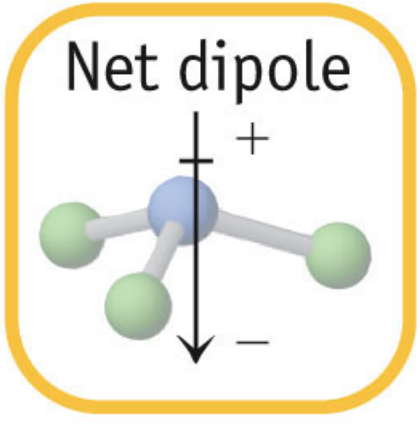
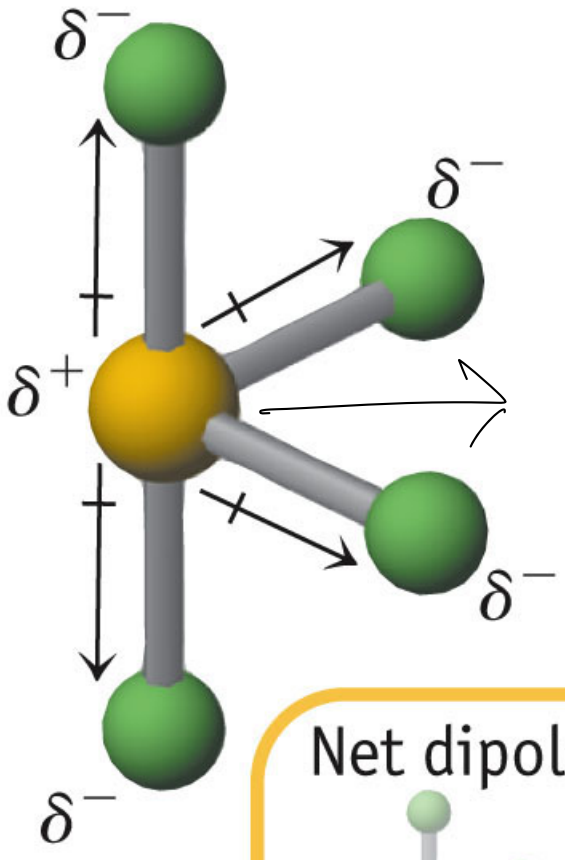


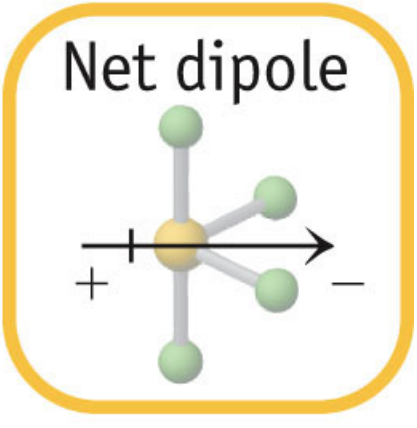
Fig. 8-14, p. 383





NF_3

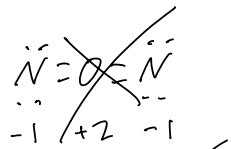


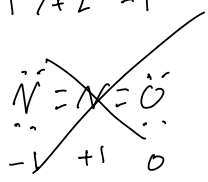


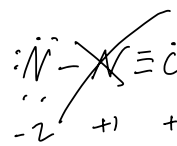
SF_4

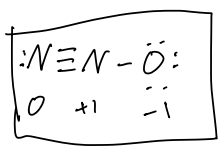
Example 8-12, p. 384

N_2O -16 e⁻









* Preferable to have any negative formal charges on more electronegative atoms

O is more electroneg than N

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

correlates w/ atom size

C–C 154 C–N 147

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

bond order → 1 = single, 2 = double, 3 = triple

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

bond length decreases as bond order increases

ΔH for breaking bond (endo thermic

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

making bonds is exothermic

N-N 163

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

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

Bonds broken

4 C-H 4×413

2 O=O 2×498

2648 kJ in

Bonds made

2 C=O 2×745 (-745)

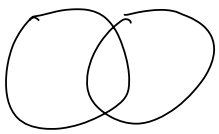
4 O-H 4×463 (-463)

3342 kJ out

Valence Bond Theory describes bonding in terms of orbital overlap

H_2 molecule - made of 2 H atoms

each H atom has $1e^-$ in a 1s orbital



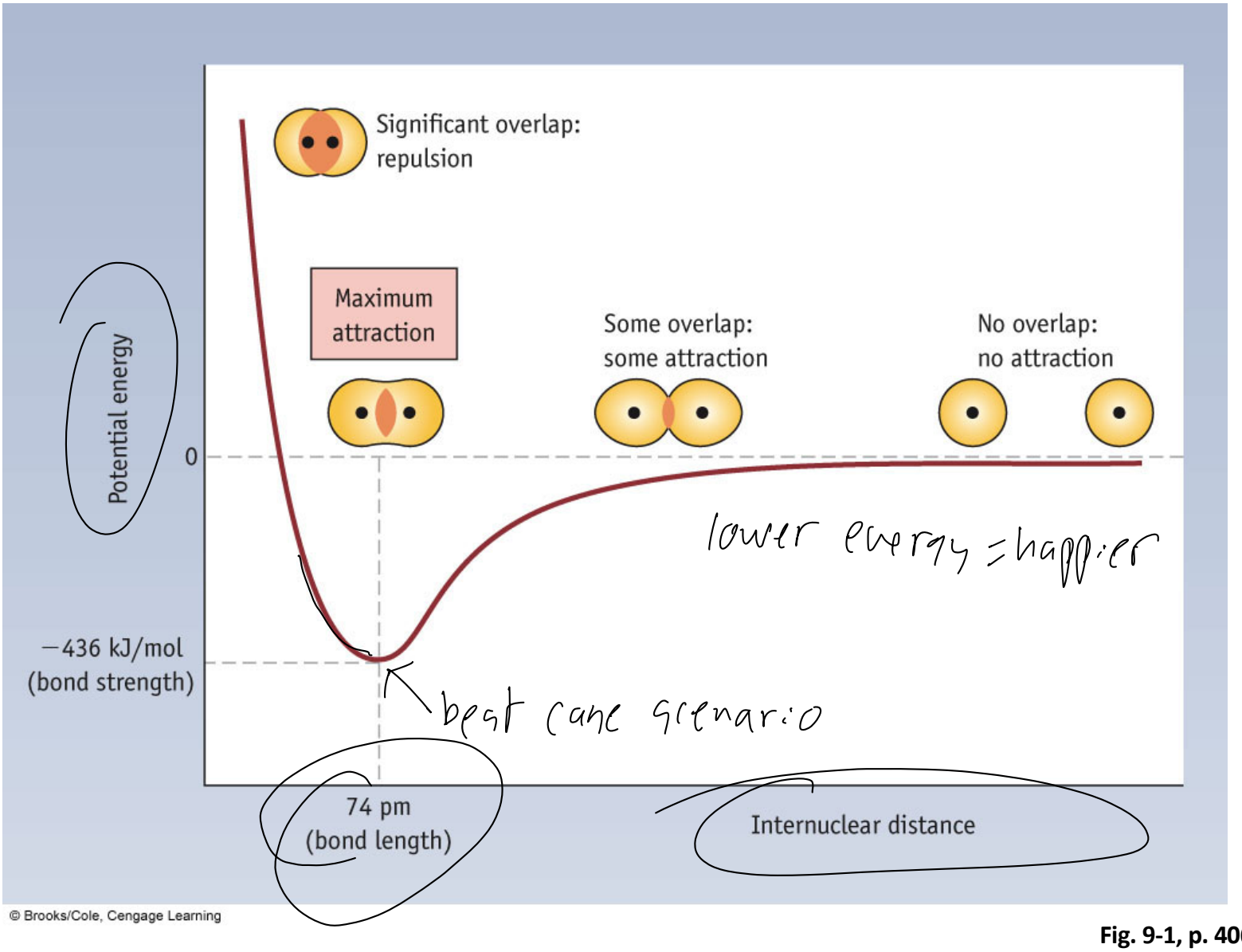


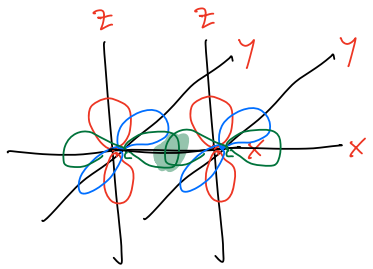
Fig. 9-1, p. 406

Cl_2 molecule - valence bond theory

:Cl–Cl: :Cl· ·Cl:

↿⇂	↿⇂	↿⇂	↿⇂	↿
3s		3p	3p	3p

unpaired 3p e⁻ used for overlap/sharing



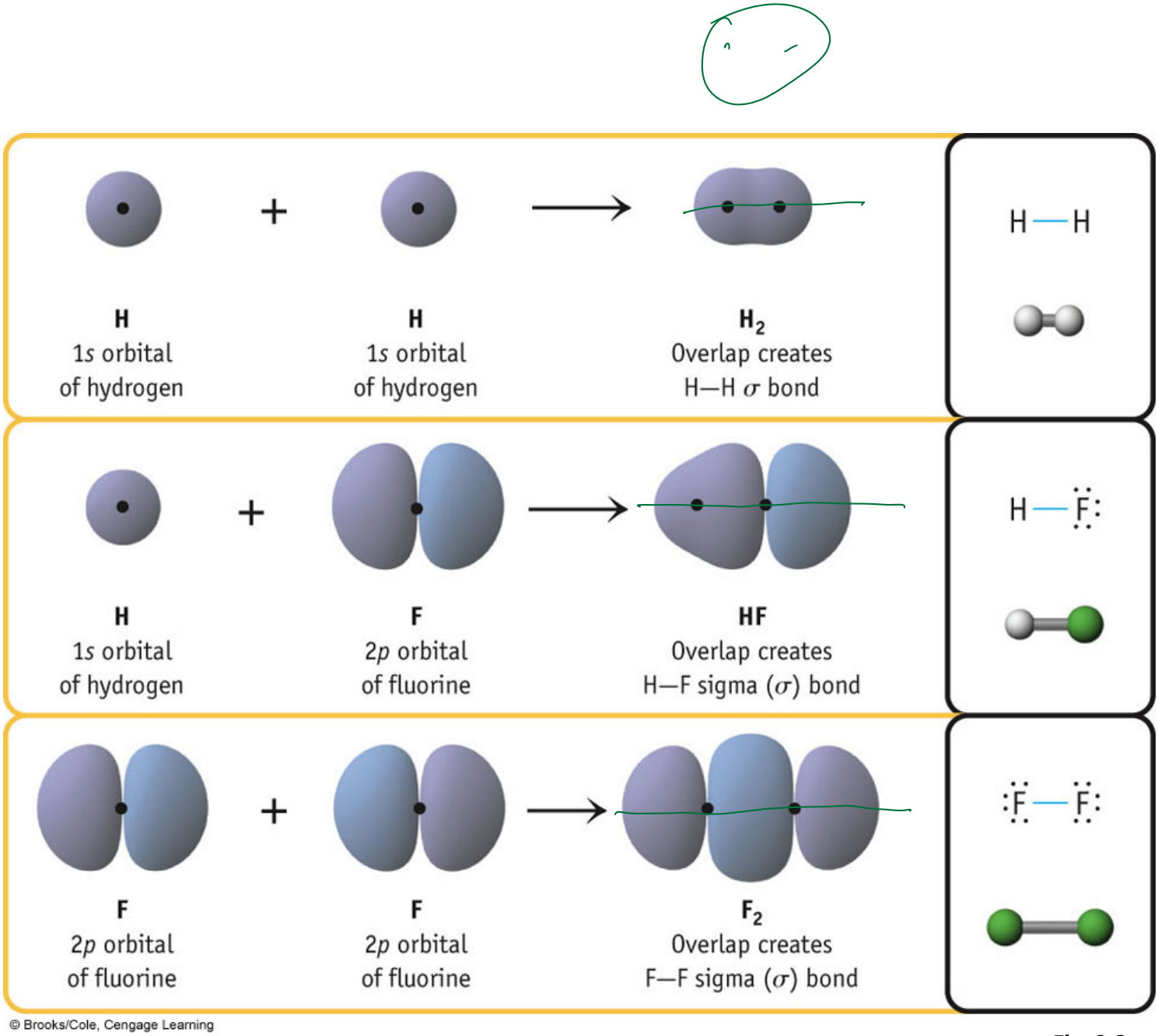


Fig. 9-2, p. 407

bond axis - line thru 2 nuclei

σ bond is when bond axis steers electron density

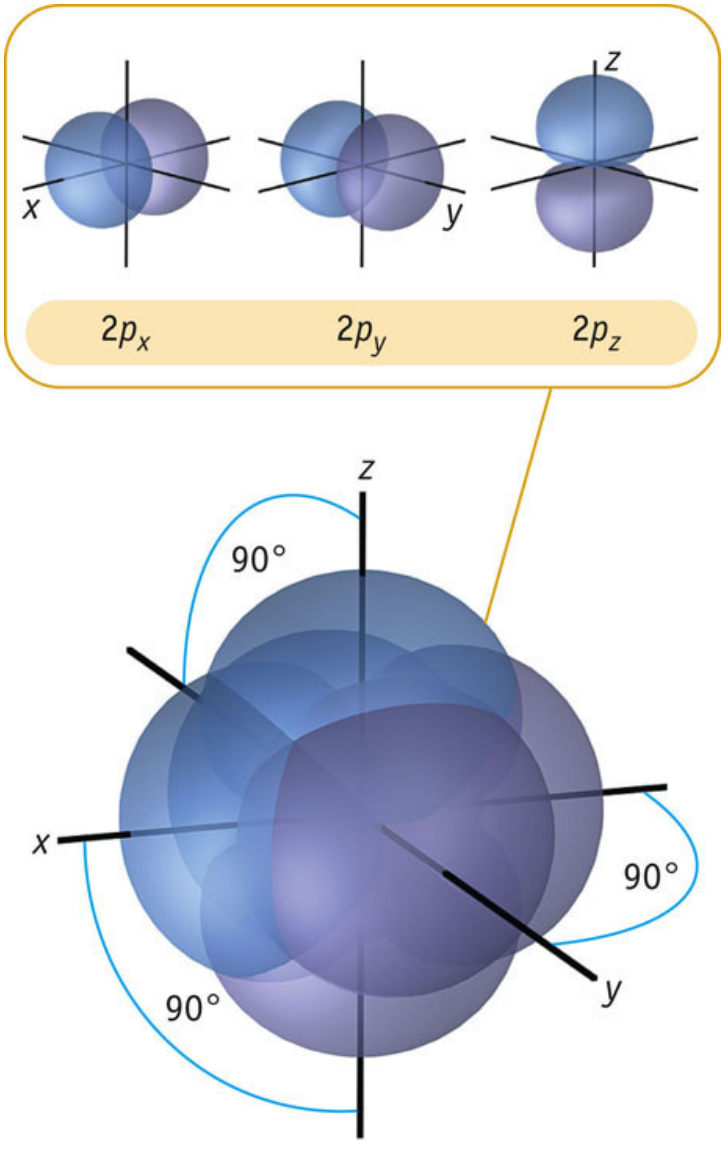
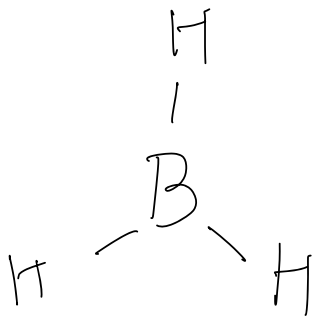


Fig. 9-3, p. 408

BH_3



trigonal planar

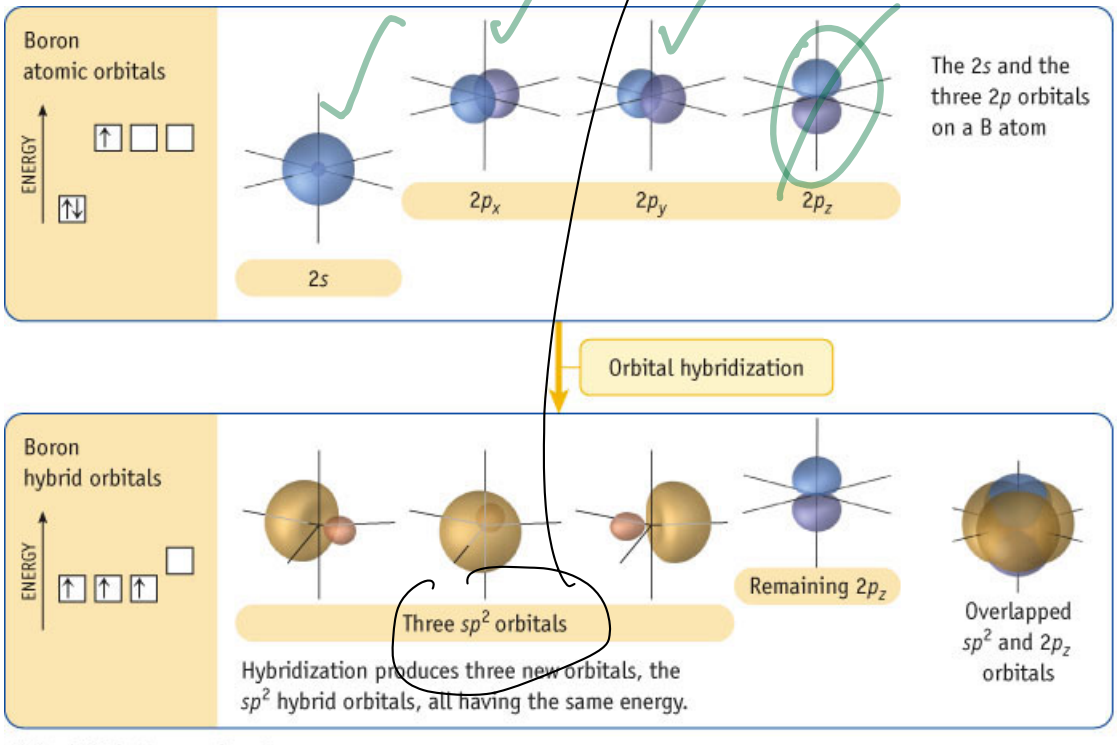
H $1s^1$

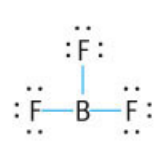
B $1s^2$ $\boxed{2s^2 2p^1}$

- need 3 B orbitals w/ e
- they need to be @ 120° to each other

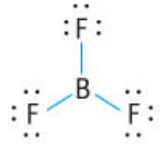
*hybridization - combine OG orbitals and make new + different orbitals w/ ~~correct~~ desired characteristics

3 new sp^2 hybrid orbitals
identical in shape, energy
oriented @ 120°

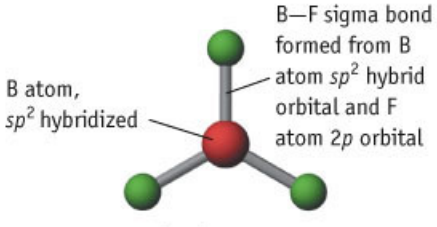




Lewis structure



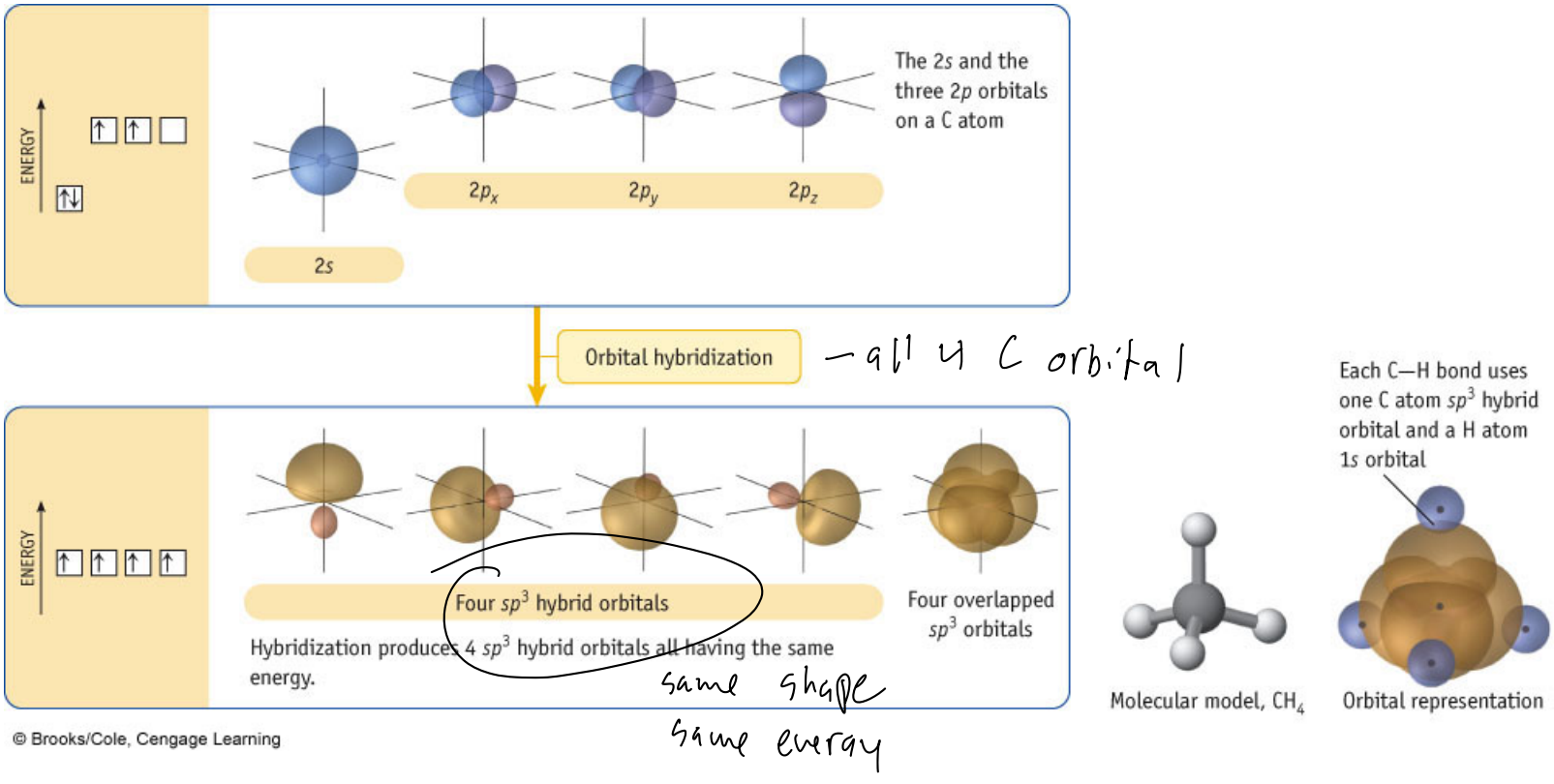
Electron-pair geometry



Molecular geometry

Fig. 9-8, p. 414

CH_4 - tetrahedral



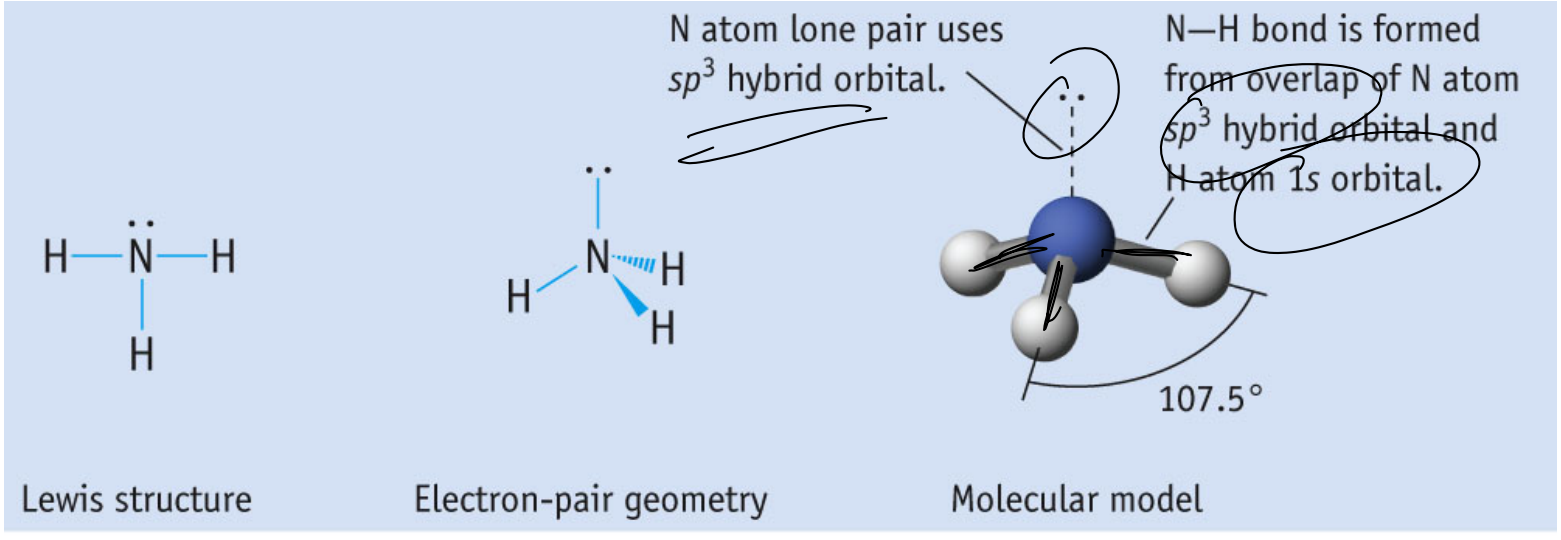
all 4 C orbital

same shape

same energy

Fig. 9-6, p. 411

tetrahedral e⁻ pair geometry → sp^3



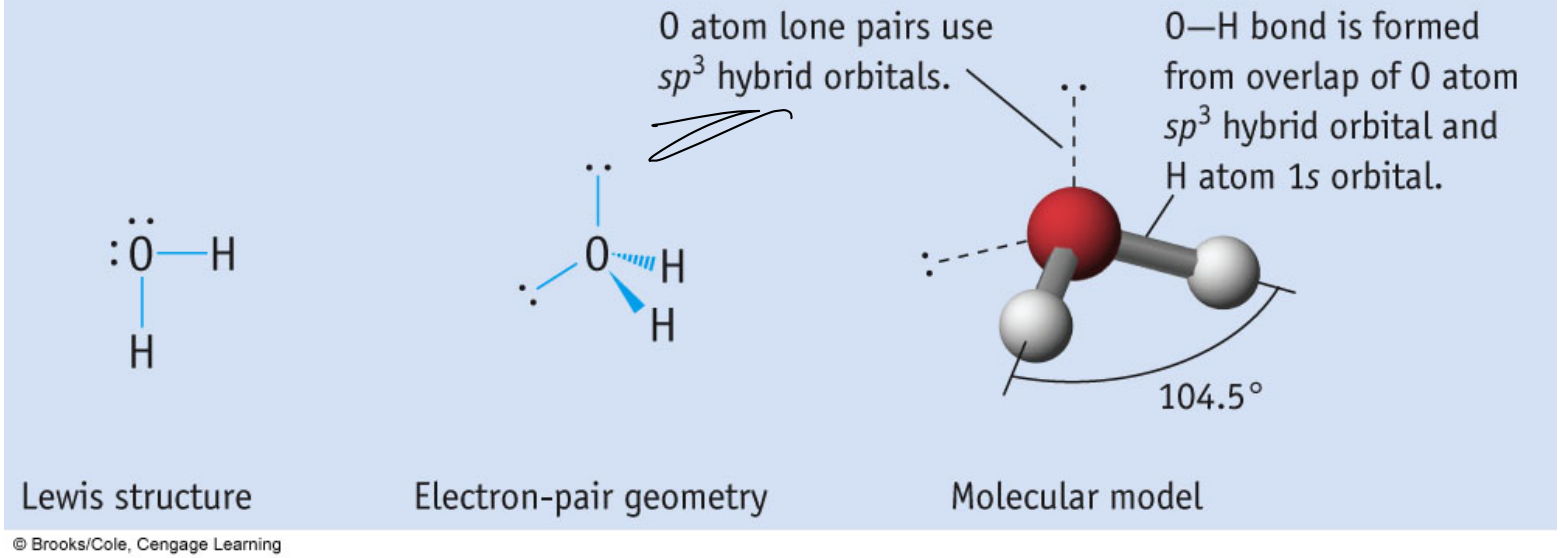


Fig. 9-7, p. 412

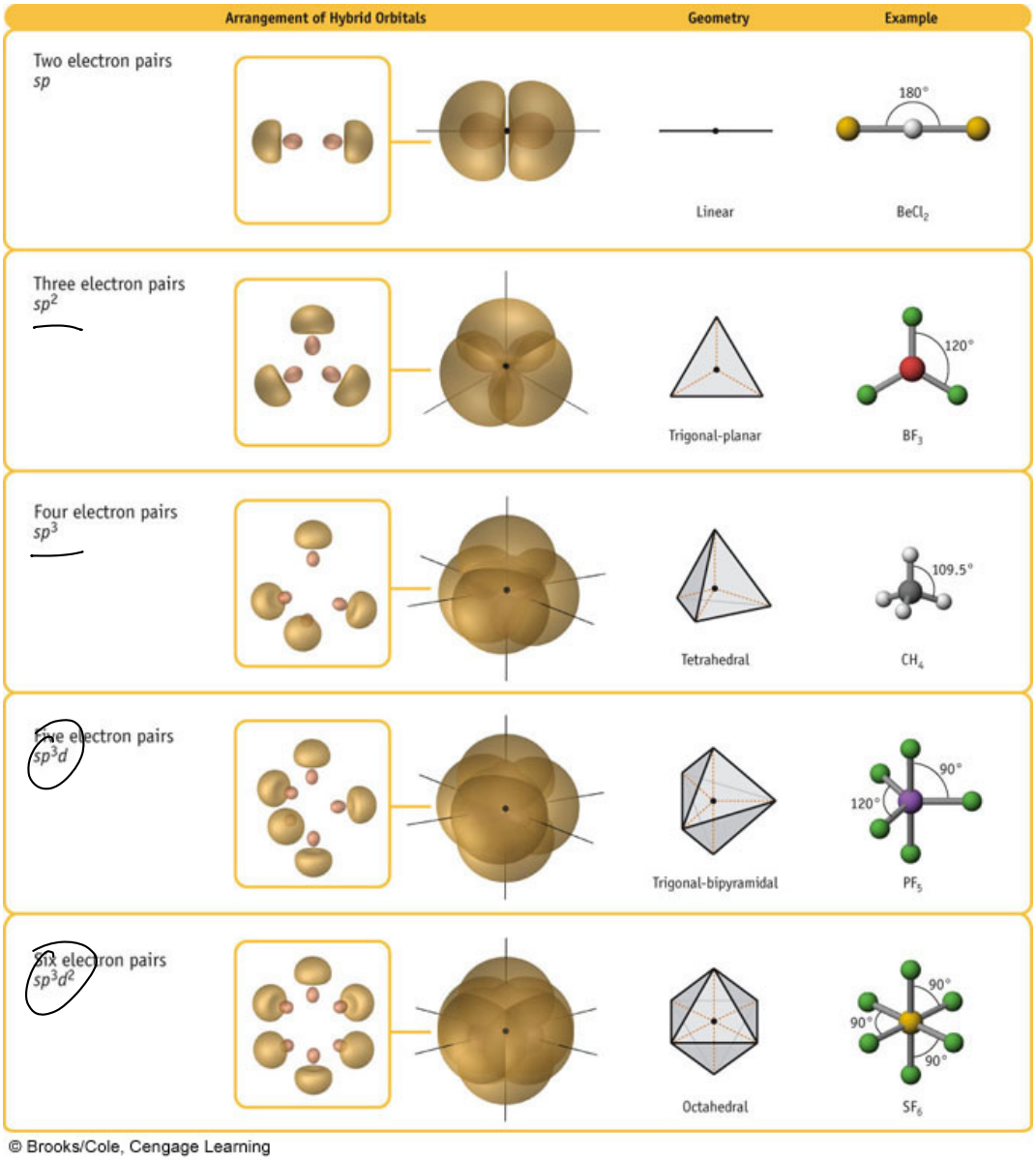
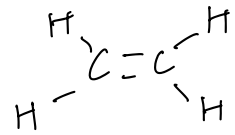


Fig. 9-5, p. 410

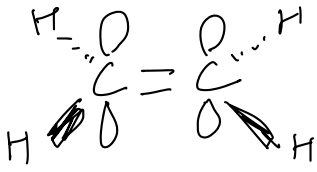
C_2H_4



trig planar @ C → sp^2 hybridized

unhybridized p orbital on each C – out of plane of molecule

carbon carbon bond



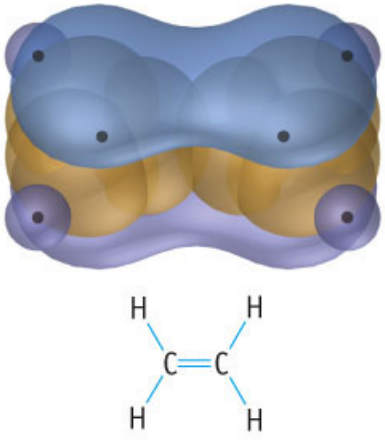
bond axis

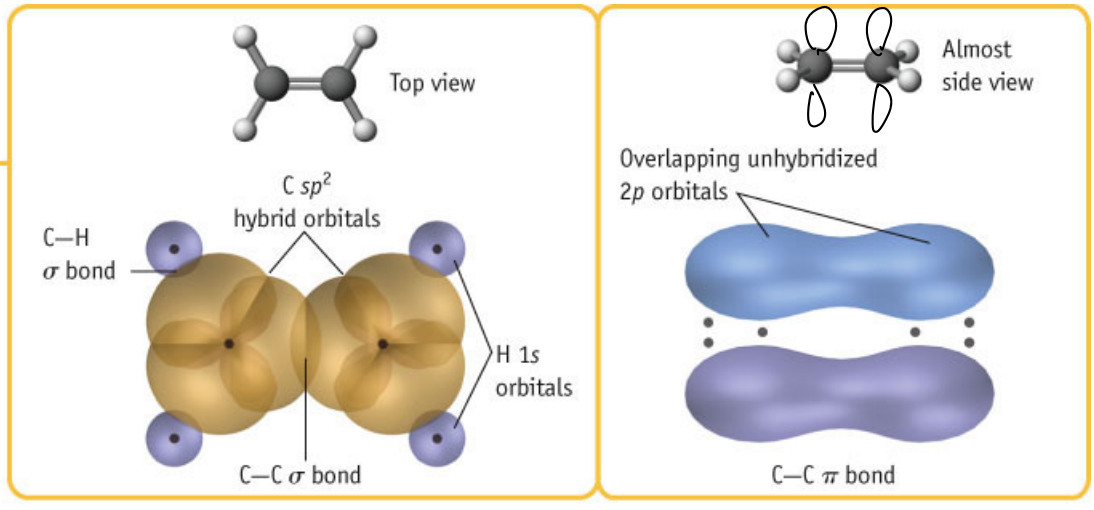
σ bond from sp^2–sp^2 overlap

π bond from unhyb p – unhyb p overlap

π bond

no electron density along bond axis





(a) Lewis structure and bonding of ethylene, C_2H_4.

(b) The C—H σ bonds are formed by overlap of C atom sp^2 hybrid orbitals with H atom 1s orbitals. The σ bond between C atoms arises from overlap of sp^2 orbitals.

(c) The carbon–carbon π bond is formed by overlap of an unhybridized 2p orbital on each atom. Note the lack of electron density along the C—C bond axis from this bond.

Fig. 9-10, p. 417

C_2H_2 $H-C\equiv C-H$

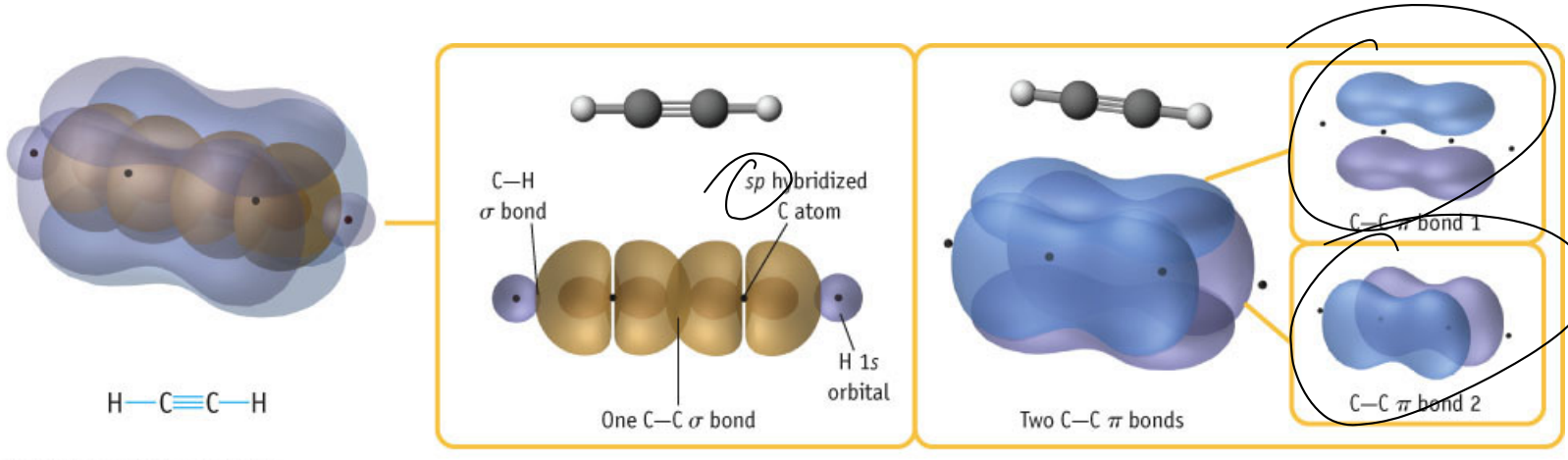
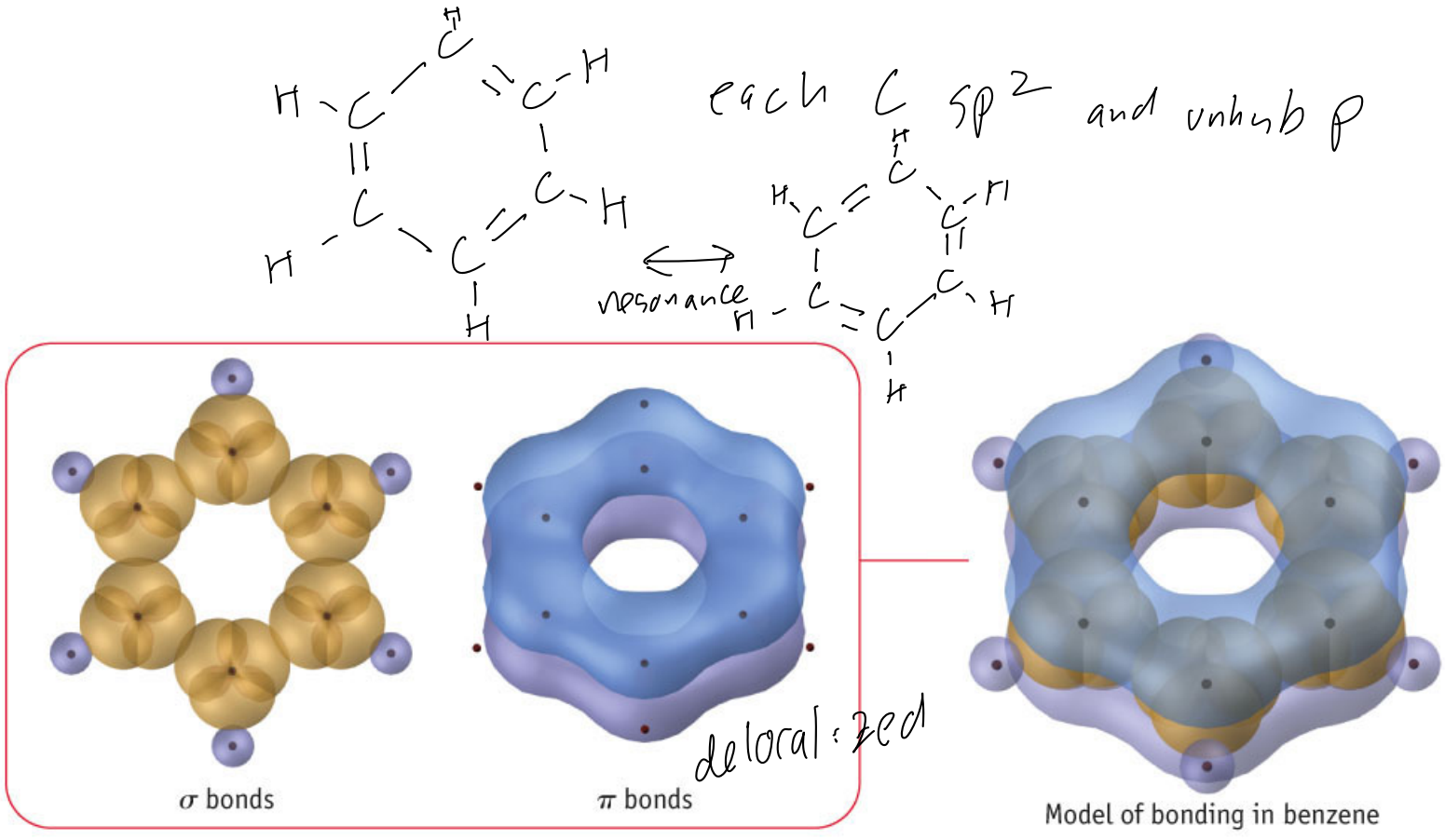


Fig. 9-12, p. 419



σ and π bonding in benzene

Fig. 9-14, p. 422

Molecular orbital theory (AO)

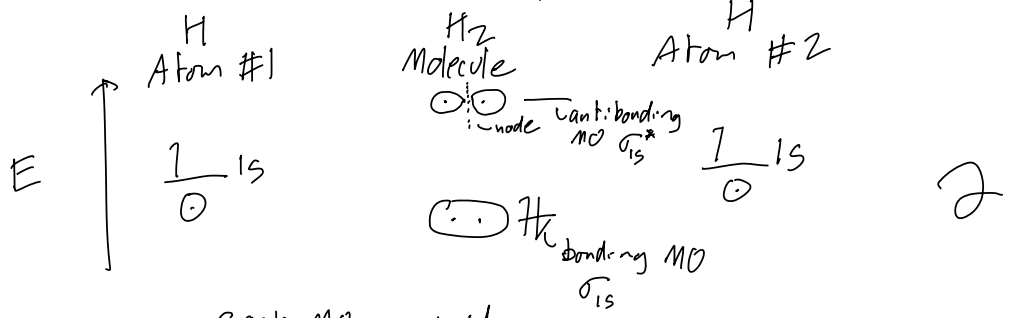
- start with regular unhyb atomic orbital
- create new orbitals that can be spread out over entire molecule → "molecular orbitals" (MO)

#AOs = #MOs

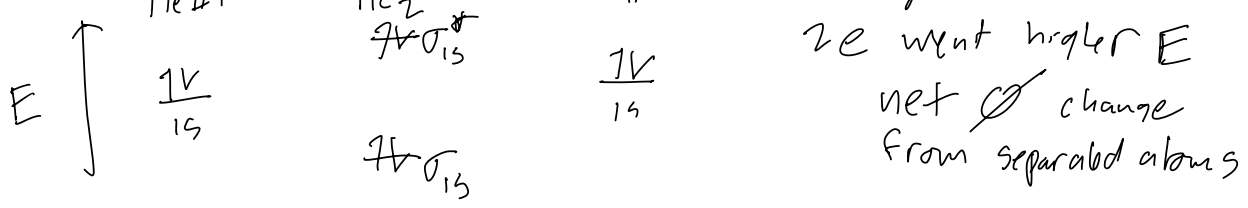
lower E than AOs

higher E than AOs

Imagine a "constructive" vs "destructive" combination



each MO can hold 2e-



2e got to lower E
2e went higher E
net ∅ change
from separated atoms

Bond order = (#e⁻ in bonding MO − #e⁻ in antibonding MO) / 2

He_2^+ → BO = 1/2

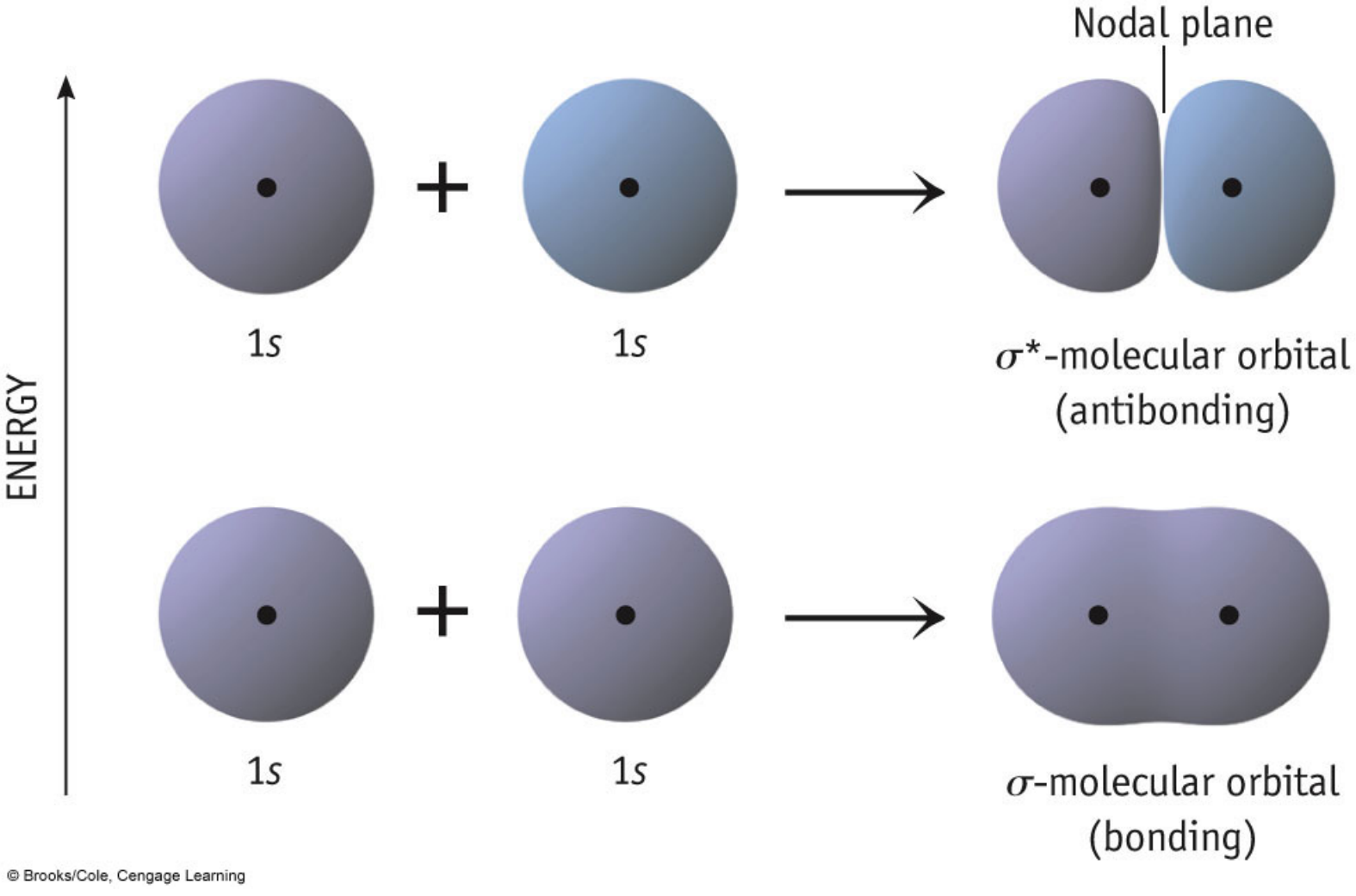


Fig. 9-16, p. 424

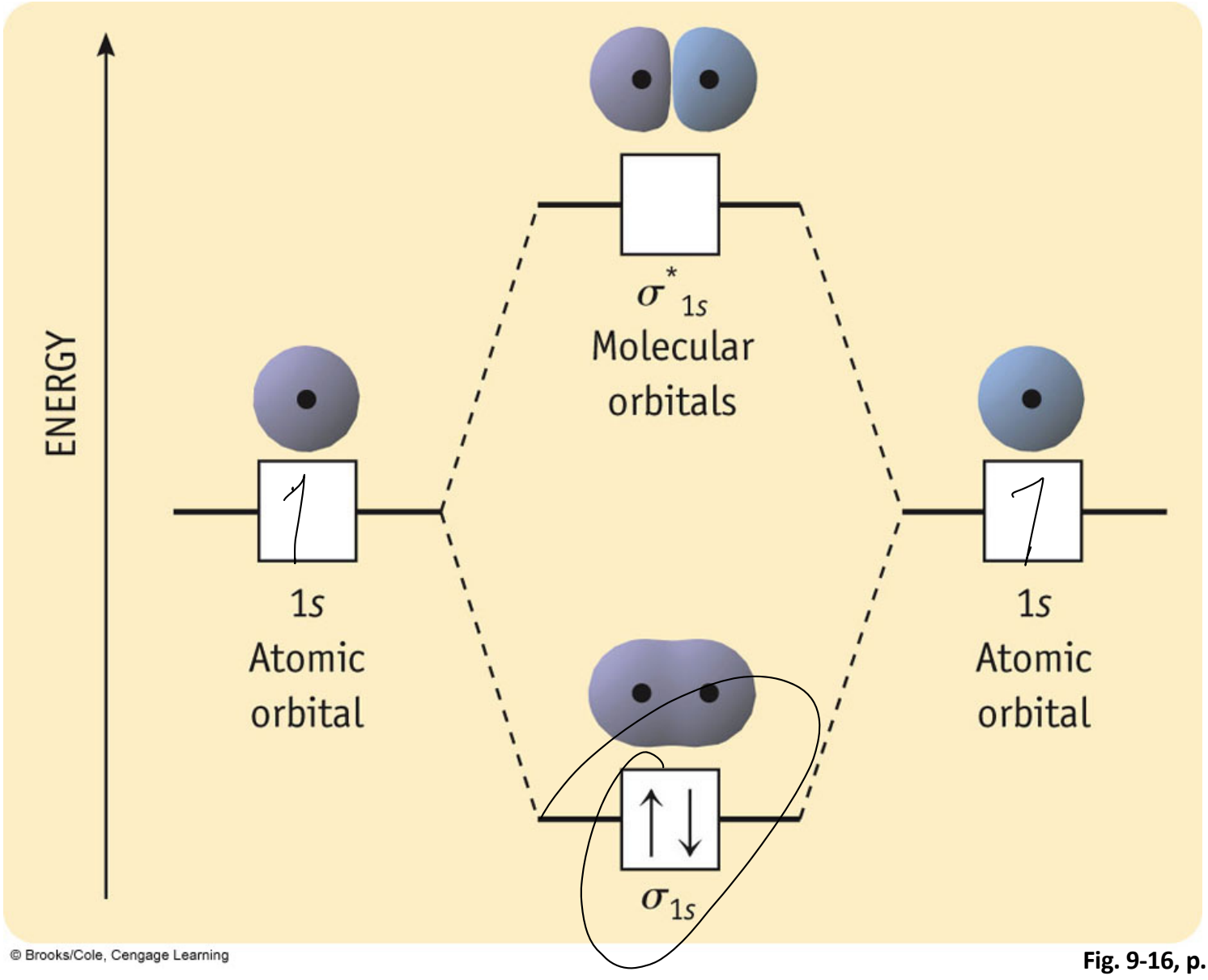


Fig. 9-16, p. 424

He2

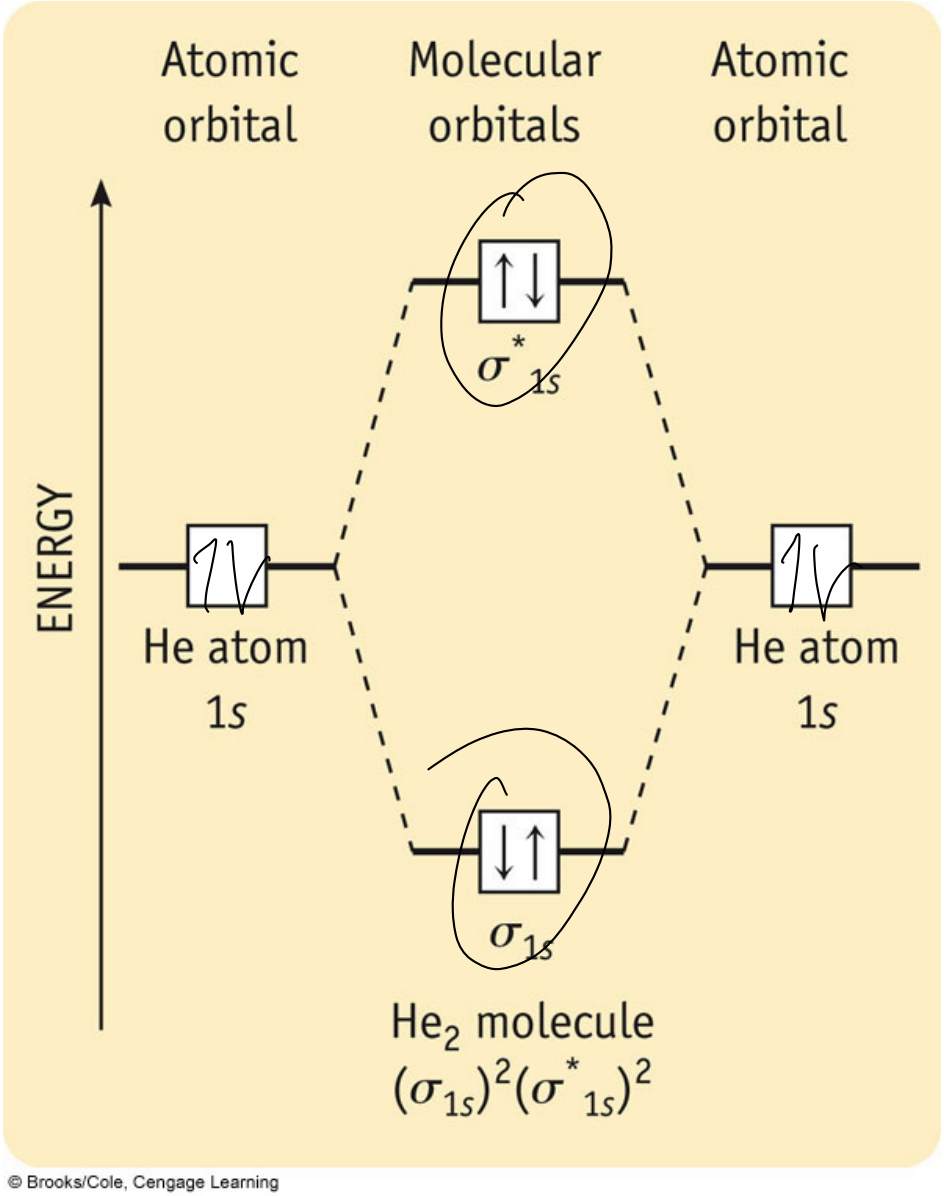


Fig. 9-17, p. 424

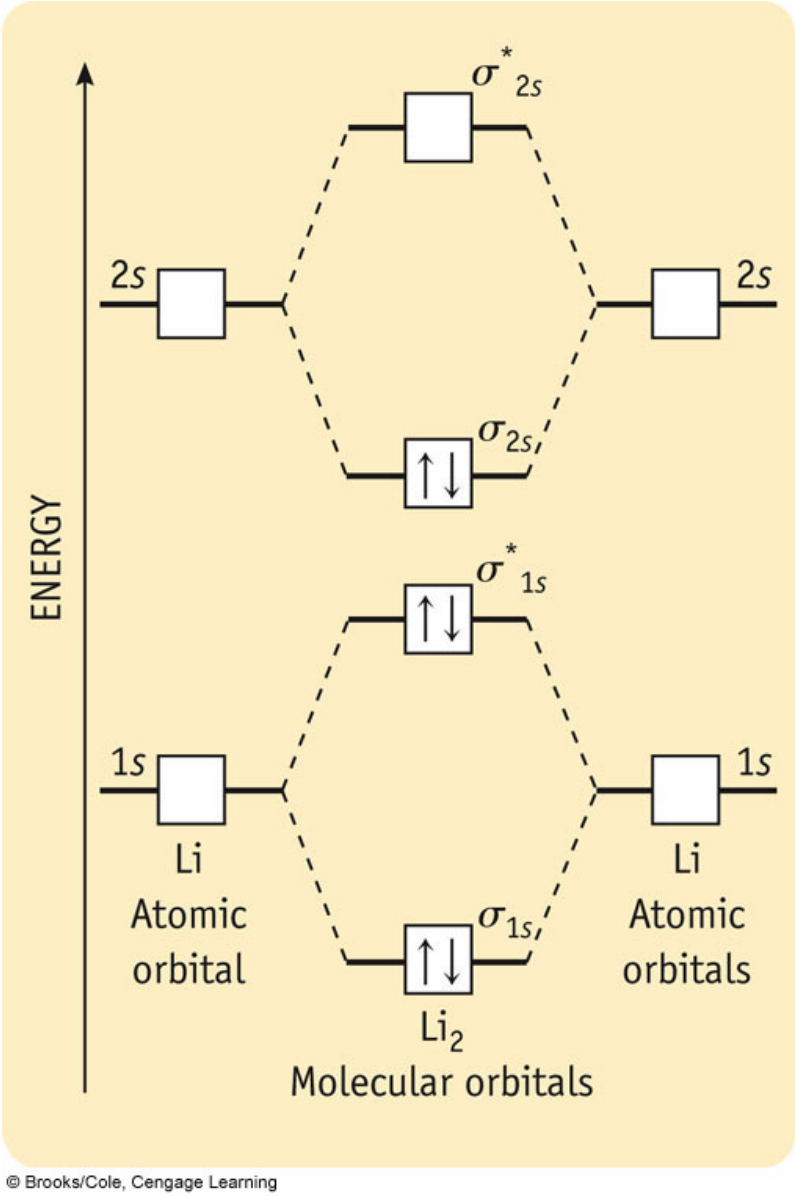
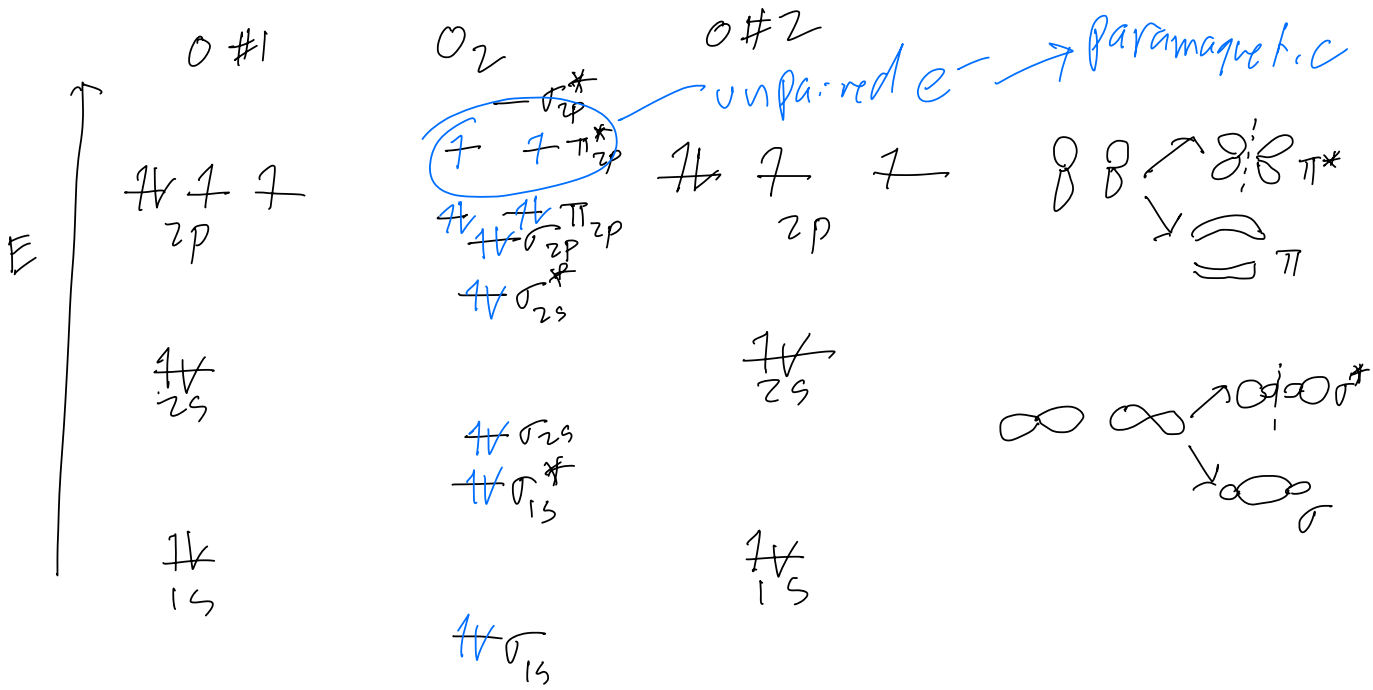


Fig. 9-18, p. 425

O_2 – MO view


O #1
O_2
O #2
unpaired e⁻ → paramagnetic
E
2p
2s
1s
σ^*_{2p}
π^*_{2p}
π_{2p}
σ_{2p}
σ^*_{2s}
σ_{2s}
σ^*_{1s}
σ_{1s}
π^*
π
σ^*
σ

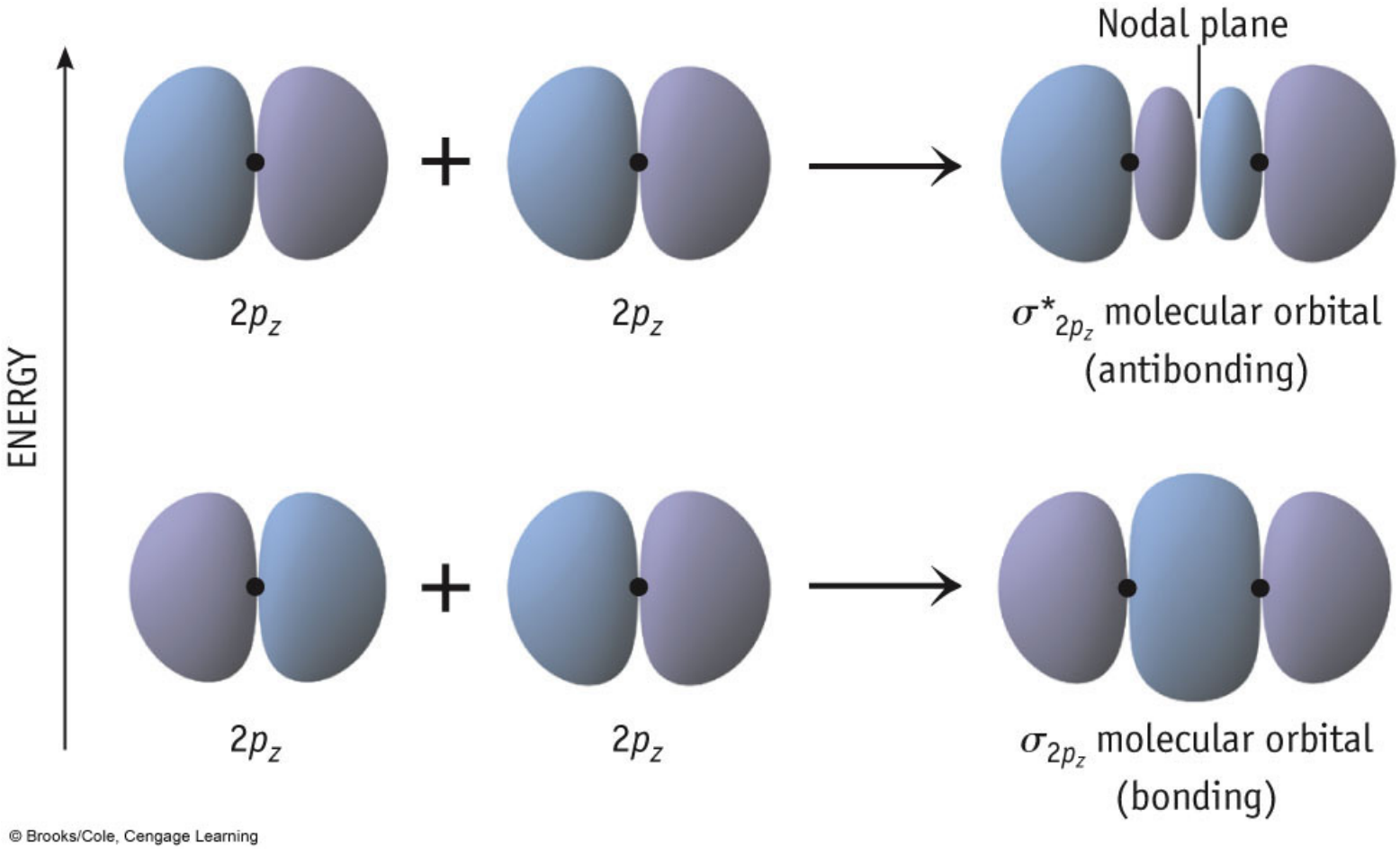


Fig. 9-19, p. 426

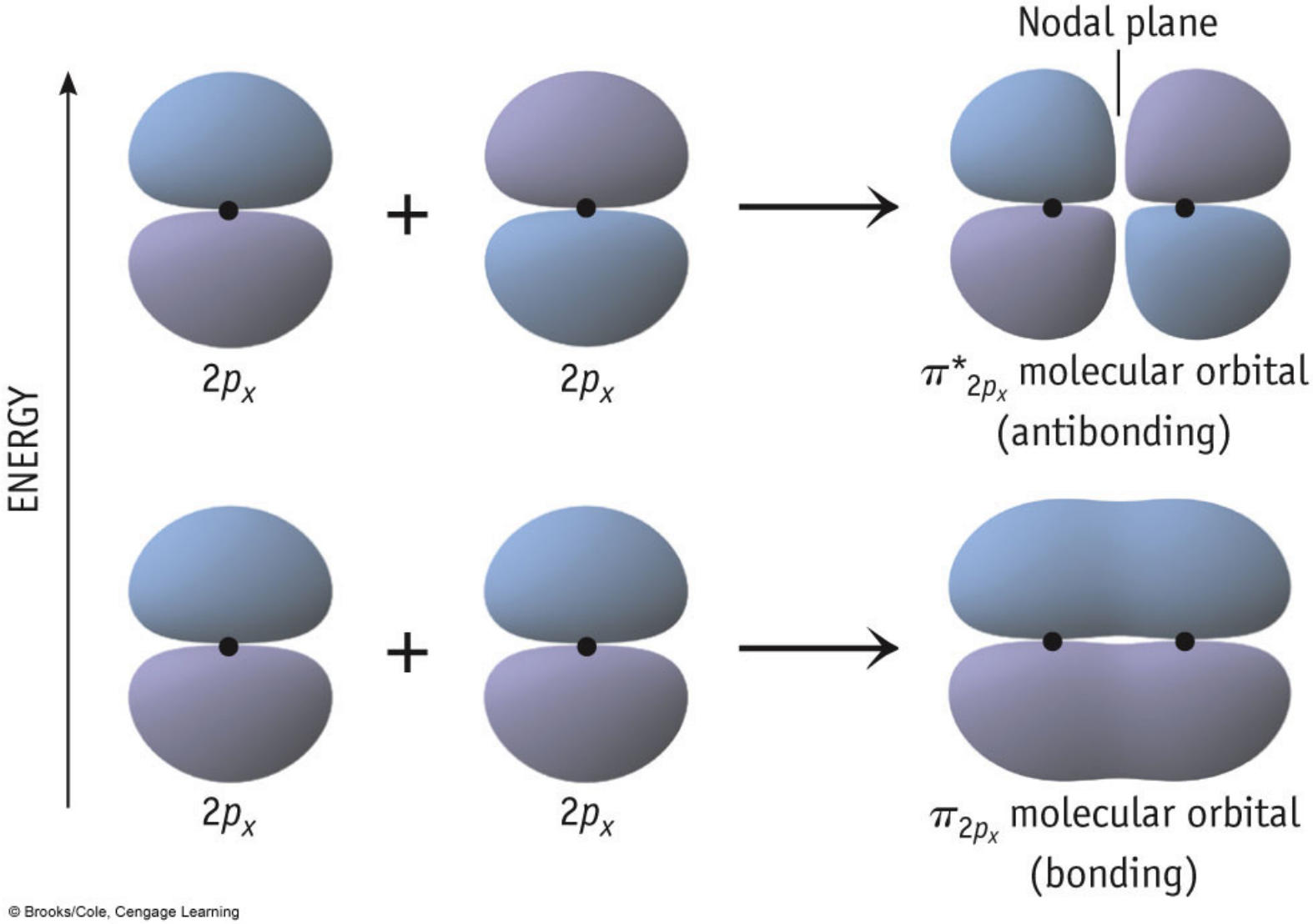


Fig. 9-20, p. 427

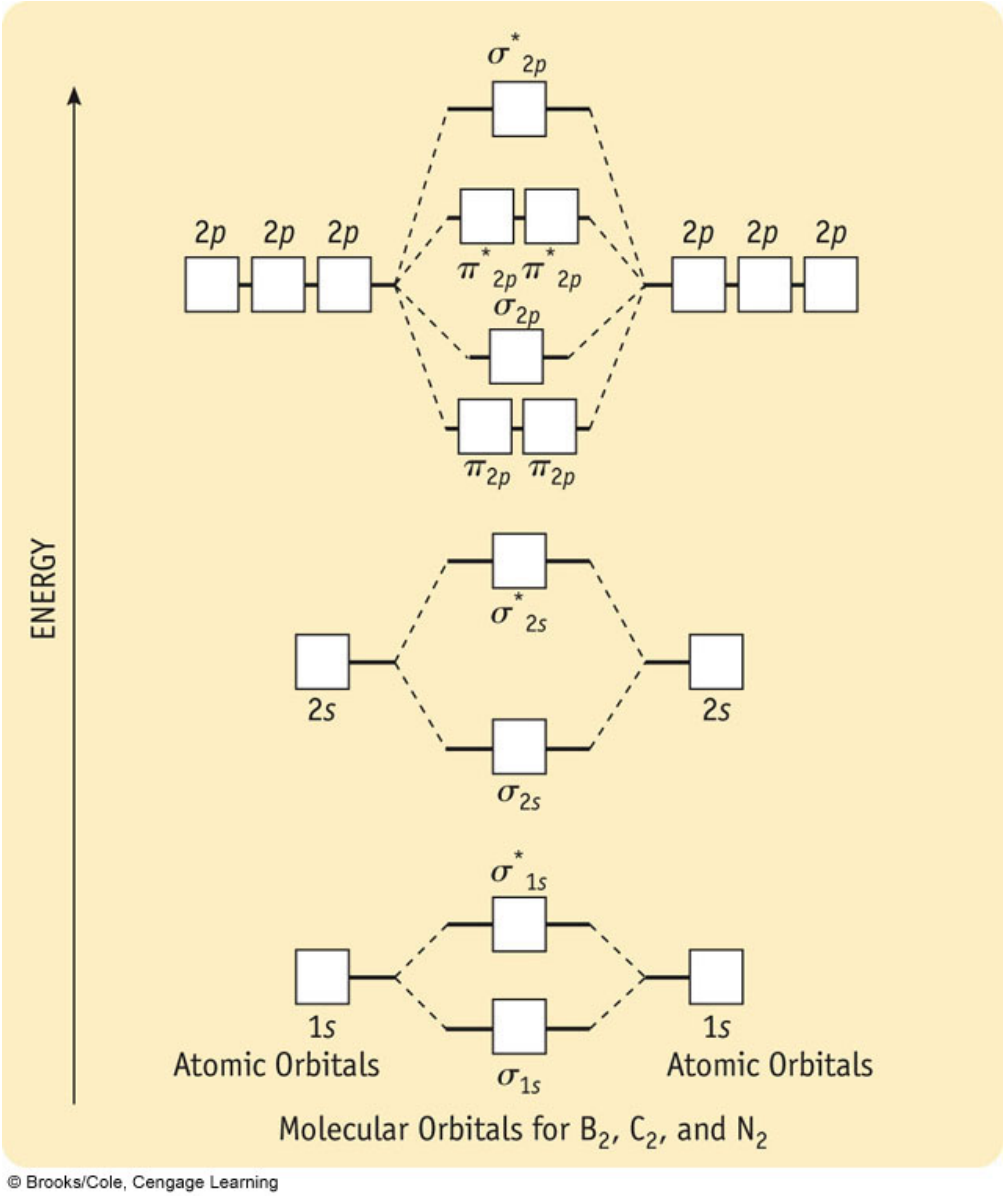


Fig. 9-21a, p. 427

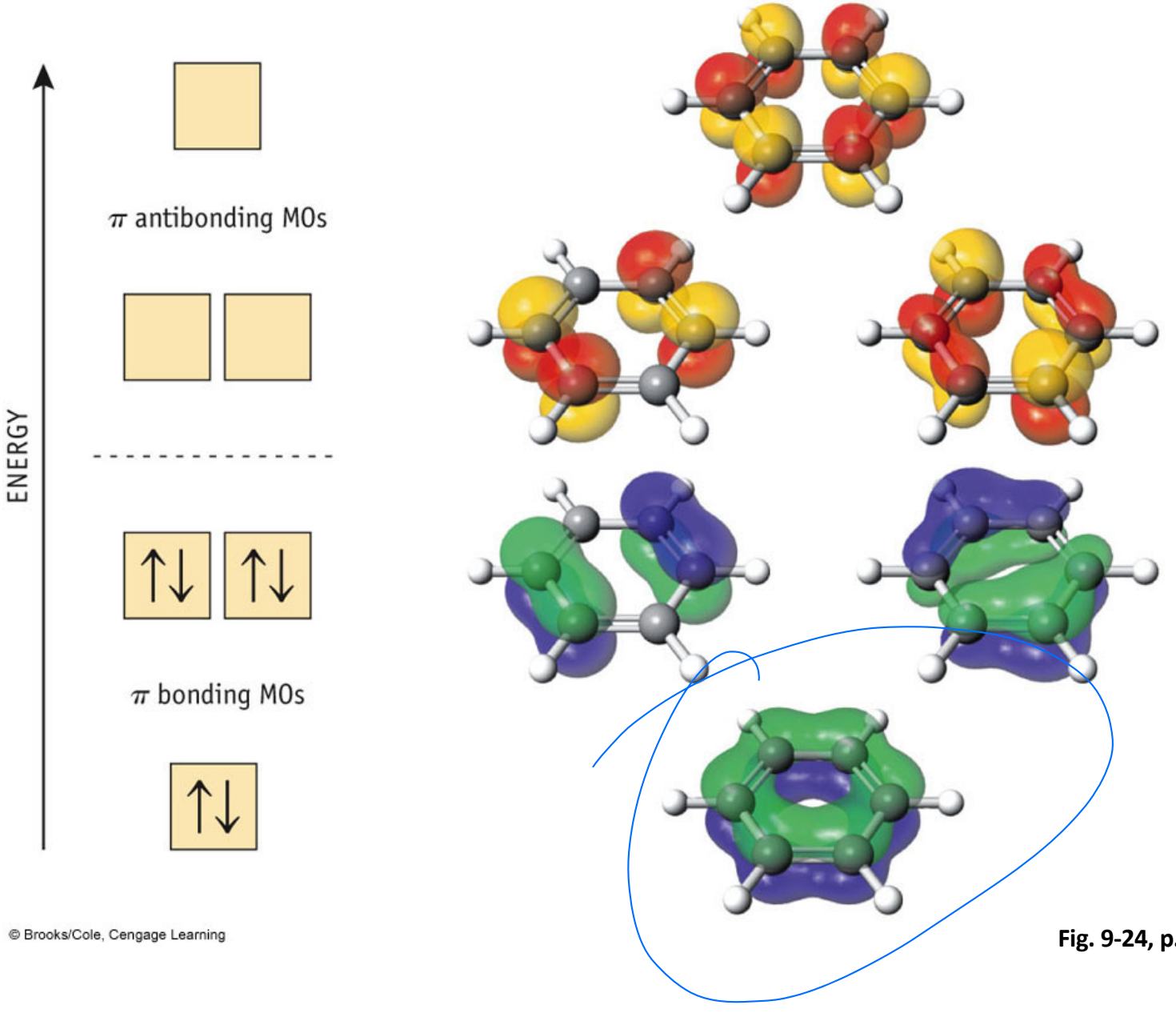


Fig. 9-24, p. 432