CH-110 Advanced General Chemistry I

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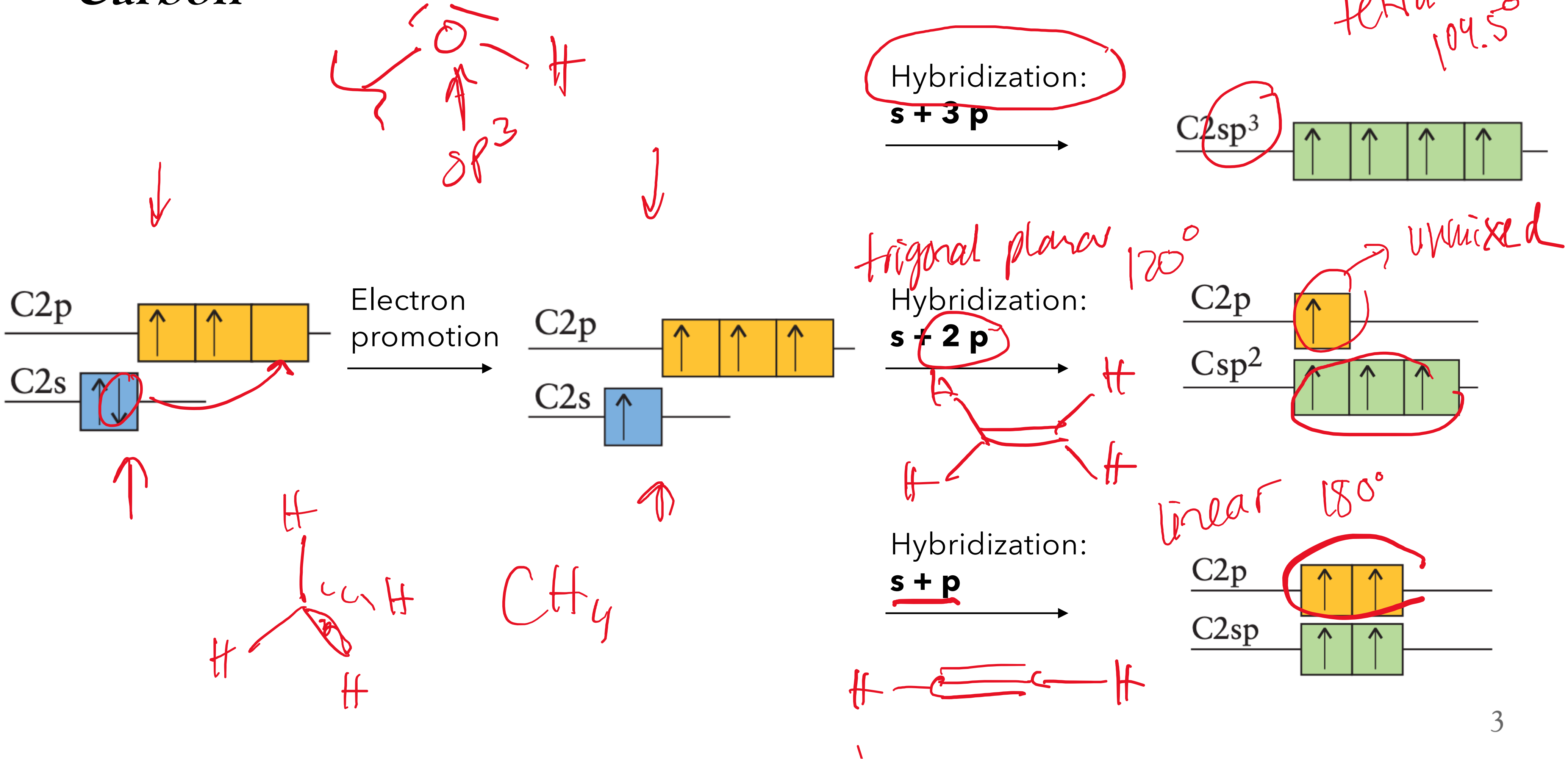
Date for review/Q&A session for the exam in January:

January 8, 14:00–16:00

Room TBD

Quick recap: Valence bond theory and hybridization

Carbon



C2p ↑ ↑

C2s ↑↓

Electron promotion

C2p ↑ ↑ ↑

C2s ↑

CH_4

tetrahedral
104.5°

Hybridization:
s + 3 p

$C2sp^3$ ↑ ↑ ↑ ↑

trigonal planar 120°

Hybridization:
s + 2 p

unmixed

C2p ↑

Csp^2 ↑ ↑ ↑

linear 180°

Hybridization:
s + p

C2p ↑ ↑

C2sp ↑ ↑

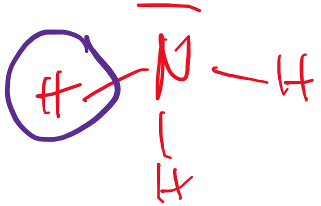
Quick recap: Valence bond theory and hybridization

Nitrogen

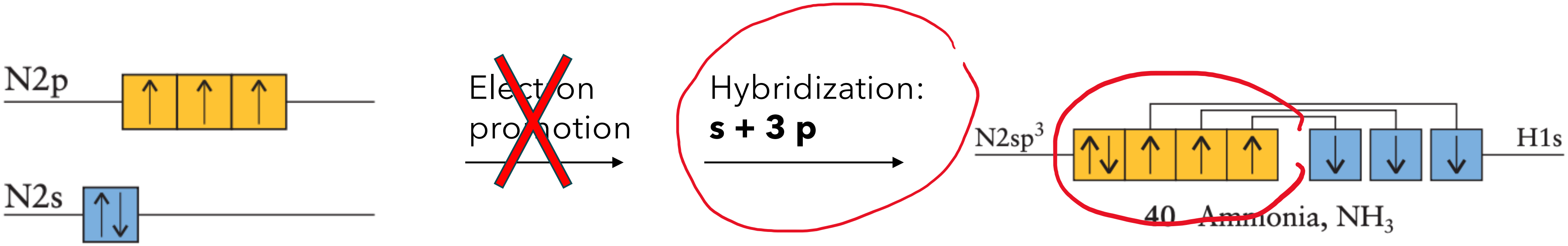
tetrahedral

↓

sp^3



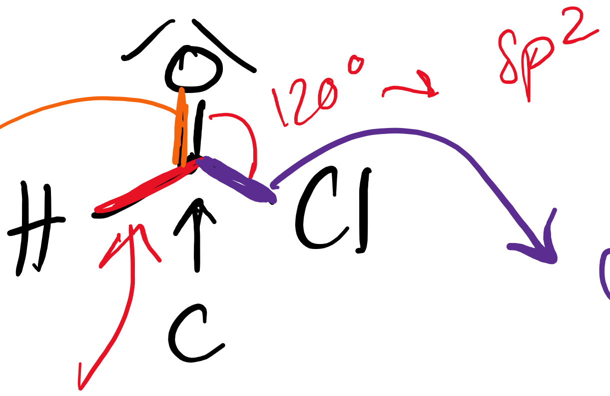
$\bar{N}H_3$



~~Electron promotion~~

Hybridization: **s + 3 p**

Final rule: We don't hybridize single-bonded terminal atoms?



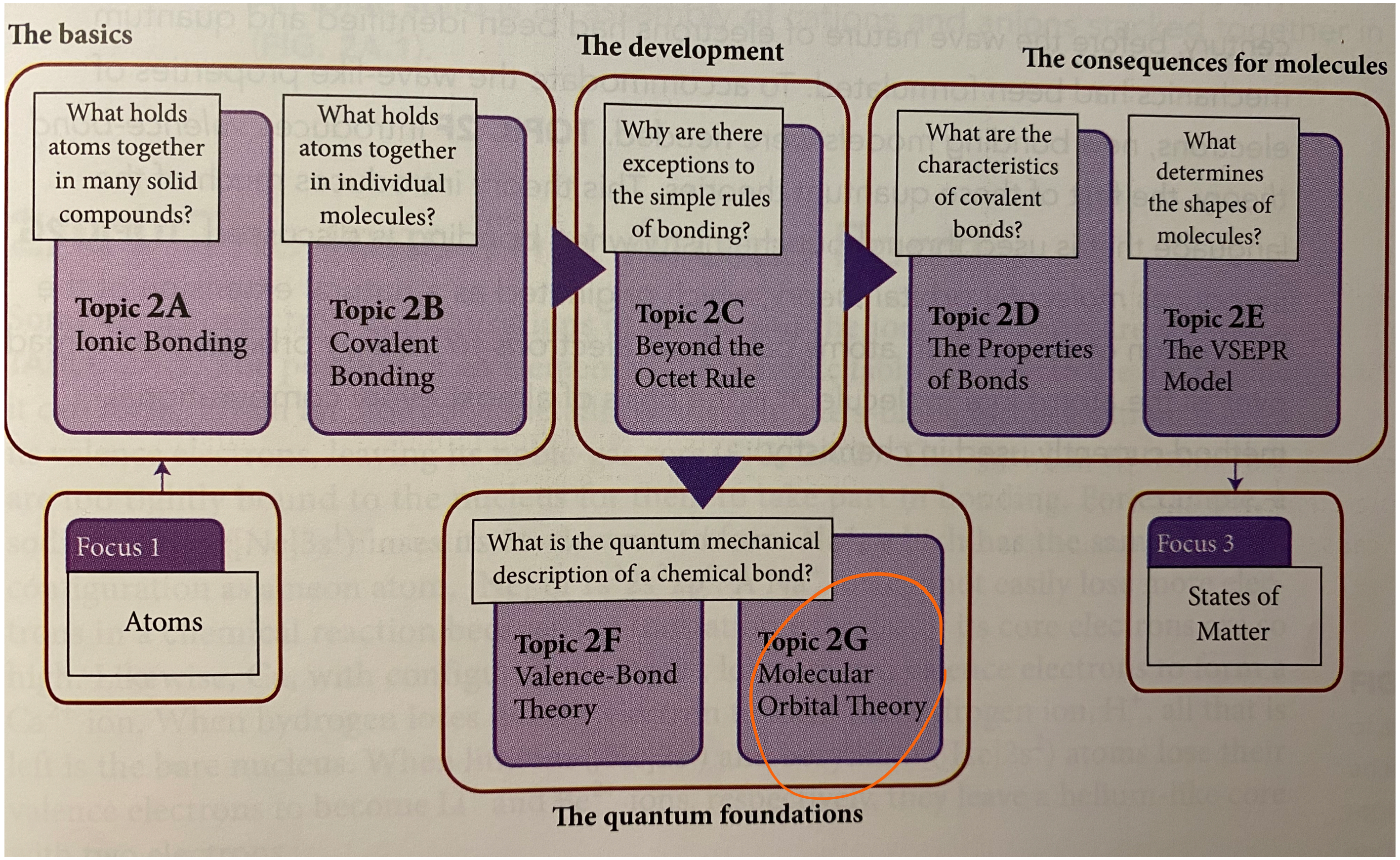
$\sigma(C2sp^2, Cl3p)$ → special rule

$\sigma(H1s, C2sp^2)$

$\sigma(C2sp^2, O2sp^2)$

$\pi(C2p, O2p)$

Overview Chapter 2 (Focus 2: Bonds Between Atoms)



The basics

What holds atoms together in many solid compounds?

Topic 2A
Ionic Bonding

What holds atoms together in individual molecules?

Topic 2B
Covalent Bonding

The development

Why are there exceptions to the simple rules of bonding?

Topic 2C
Beyond the Octet Rule

The consequences for molecules

What are the characteristics of covalent bonds?

Topic 2D
The Properties of Bonds

What determines the shapes of molecules?

Topic 2E
The VSEPR Model

Focus 1
Atoms

What is the quantum mechanical description of a chemical bond?

Topic 2F
Valence-Bond Theory

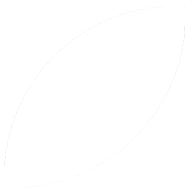
Topic 2G
Molecular Orbital Theory

The quantum foundations

Focus 3
States of Matter

Molecular Orbital Theory

Topic 2G



Topic 2G.1 Molecular orbitals
Topic 2G.2 Electron configurations of diatomic molecules
Topic 2G.3 Bonding in heteronuclear diatomic molecules
~~Topic 2G.4 Orbitals in polyatomic molecules~~
Topic 2G.5 A comparison of bonding models

Diatomic molecules

WHY DO YOU NEED TO KNOW THIS MATERIAL?

- MO theory is the most common quantum mechanical approach used to describe electronic structure of molecules.
- Essential for understanding the properties of individual molecules and modern materials.

WHAT DO YOU NEED TO KNOW ALREADY?

- Atomic orbitals (Topic 1D)
- Born interpretation of wavefunction (Topic 1C)
- Building-up principle (Topic 1E)
- Electronegativity (Topics 1F and 2D)

2G.1 Molecular orbitals

Valence bond theory deficiencies

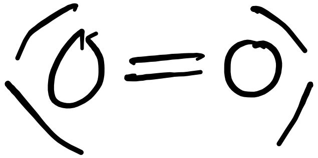


Figure 2G.10 (new book)

We know that oxygen is paramagnetic:

https://www.youtube.com/watch?v=Lt4P6ctf06Q

- Attracted to magnetic fields
- O_2 must have unpaired electrons
- Paired electrons: spins are opposite, magnetic moments cancel each other out.

→ This is not explained by VB theory.

A new theory is needed.

2G.1 Molecular orbitals

Molecular orbital (MO) theory vs. valence bond (VB) theory

MO theory

- **Addresses some limitation**s of VB theory
- Provides a deeper understanding of electron-pair bonds
- Easier to calculate computationally than VB theory
- Is used to calculate **energies** of orbitals

In MO theory: electrons occupy MOs that are **delocalized** over entire molecule

In VB theory: electrons are **localized** between the two atoms

2G.1 Molecular orbitals

Molecular orbital (MO) theory vs. valence bond (VB) theory

Aspect	VB Theory	MO Theory
Wave Function Mixing	Combines atomic orbitals into hybrid orbitals	Combines atomic orbitals into molecular orbitals
Localization	Localized between specific atoms	Delocalized over the entire molecule
Bonding Description	Overlap of orbitals, hybridization	Bonding and anti-bonding orbitals, delocalized π bonds

2G.1 Molecular orbitals

Linear combinations of atomic orbitals

- **Making bonds = mixing wavefunctions of atoms**
- The technical term for adding together wavefunctions is «forming a linear combination».

$$\Psi = \Psi_{A1s} + \Psi_{B1s}$$

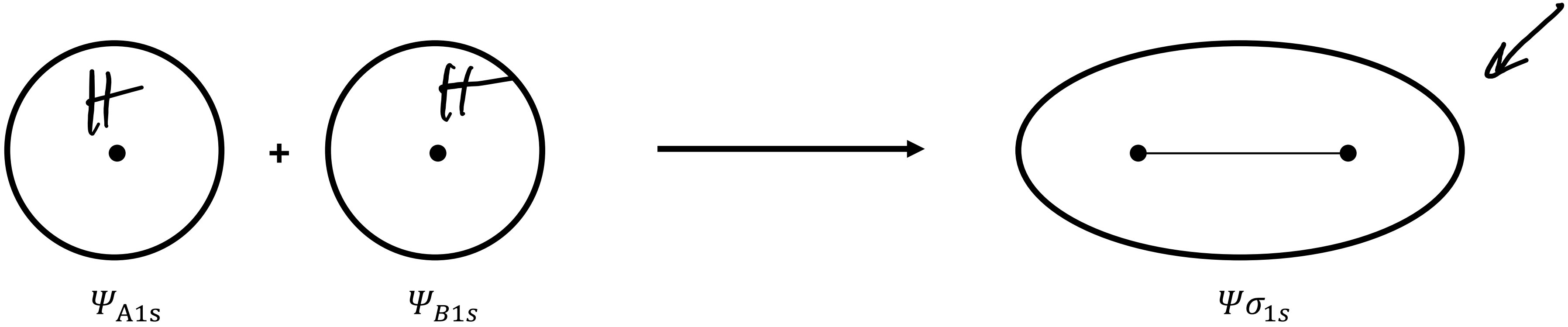
- This equation represents a **linear combination of atomic orbitals (LCAO).**
- A molecular orbital formed from a linear combination of atomic orbitals on different atoms is called an **LCAO-MO**.
- **The combination of *N* atomic orbitals results in the formation of *N* molecular orbitals.**

2G.1 Molecular orbitals

Bonding orbitals

H_2

Bonding orbitals arise from a linear combination of atomic orbitals (LCAO) (**constructive** interference).



Atomic orbitals

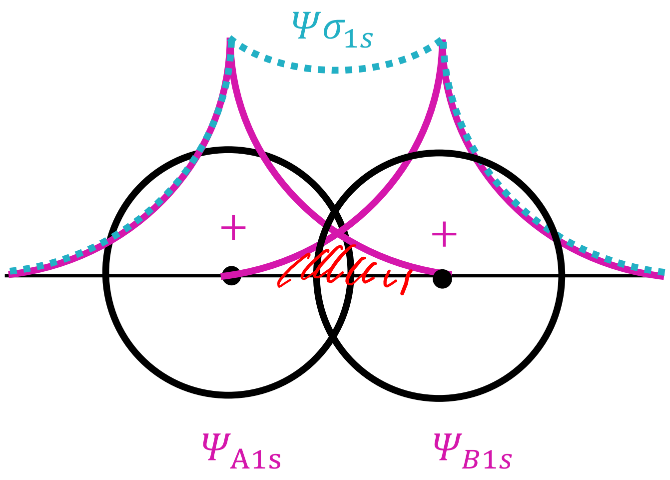
Molecular orbital

σ: cylindrically symmetric about the bond axis. No nodal plane along the bond axis.

Bonding orbital: $\Psi_{A1s} + \Psi_{B1s} = \Psi_{\sigma(1s)}$

2G.1 Molecular orbitals

Bonding orbitals



When waves interfere **constructively**, the amplitude **increases** where they overlap.

Increased amplitude in the internuclear region translates to an enhanced probability density (Ψ^2) between the nuclei.

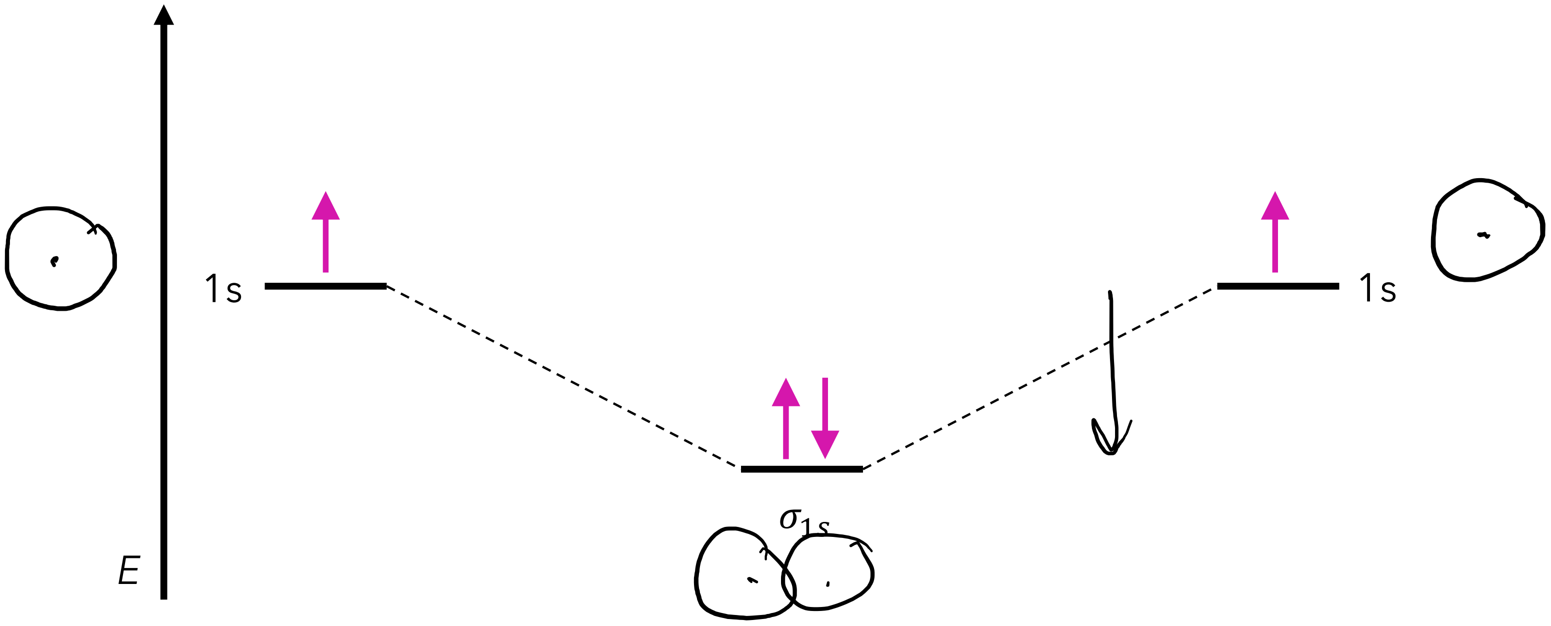
An electron in a **bonding MO** will be attracted to BOTH nuclei, and will have a **lower energy** compared to an atomic orbital for a single nucleus.

2G.1 Molecular orbitals

The energy of a bonding orbital is decreased compared to the atomic orbitals!

picture incomplete!

Molecule: H_2

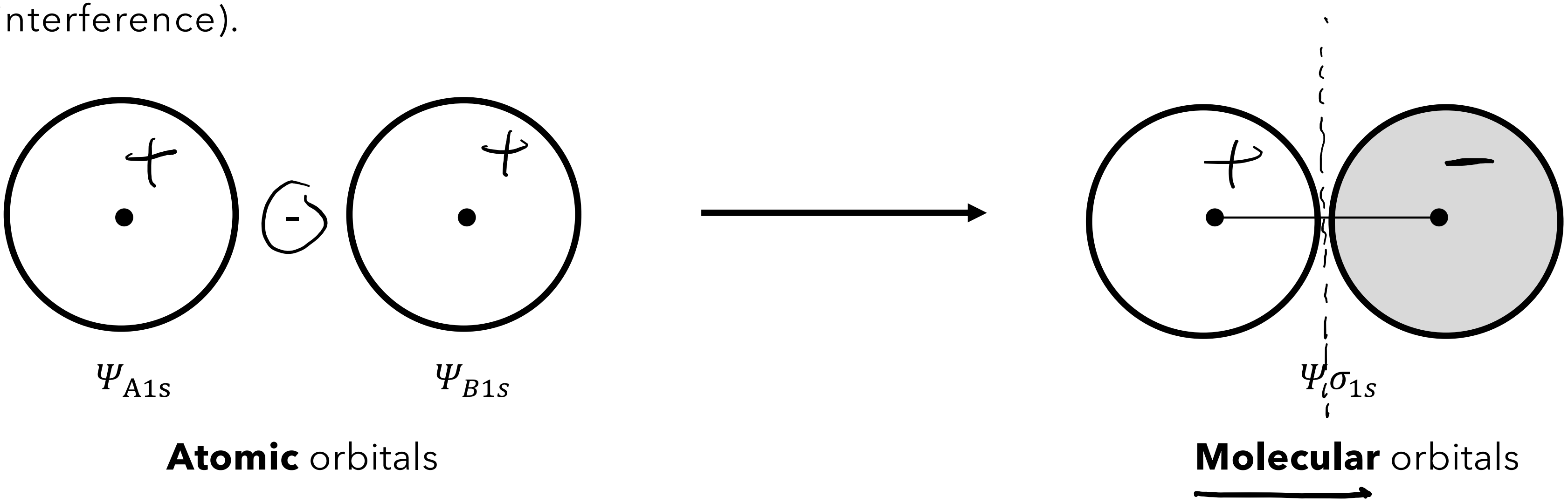


For H_2, when its two electrons both occupy the bonding orbital, the molecule is **more** stable.

2G.1 Molecular orbitals

Antibonding orbitals

Bonding orbitals arise from a linear combination of atomic orbitals (LCAO) (**destructive** interference).



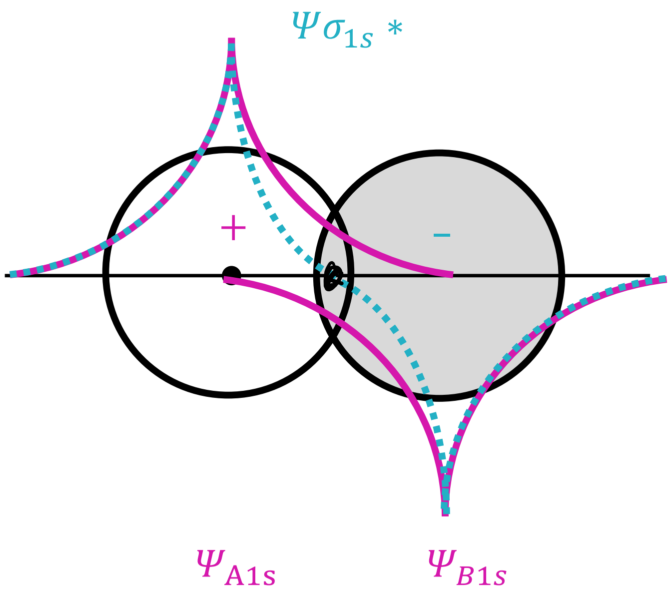
Atomic orbitals

Molecular orbitals

Antibonding orbital: $\Psi_{A1s} - \Psi_{B1s} = \Psi_{\sigma(1s^*)}$

2G.1 Molecular orbitals

Antibonding orbitals



When waves interfere **destructively**, the amplitude **decreases** where they overlap.

Decreased amplitude in the internuclear region translates to a diminshed probability density (Ψ^2) between the nuclei and a **node** between the two nuclei.

An electron in an **antibonding MO** will be essentially excluded from the internuclear region, and thus have a **higher energy** compared to an atomic orbital for a single nucleus.

2G.1 Molecular orbitals

Drawing MO diagrams

MO theory adopts the approach used for **predicting the electron configurations of many-electron atoms** (Topic 1E) but applies it to **molecular orbitals** rather than atomic orbitals.

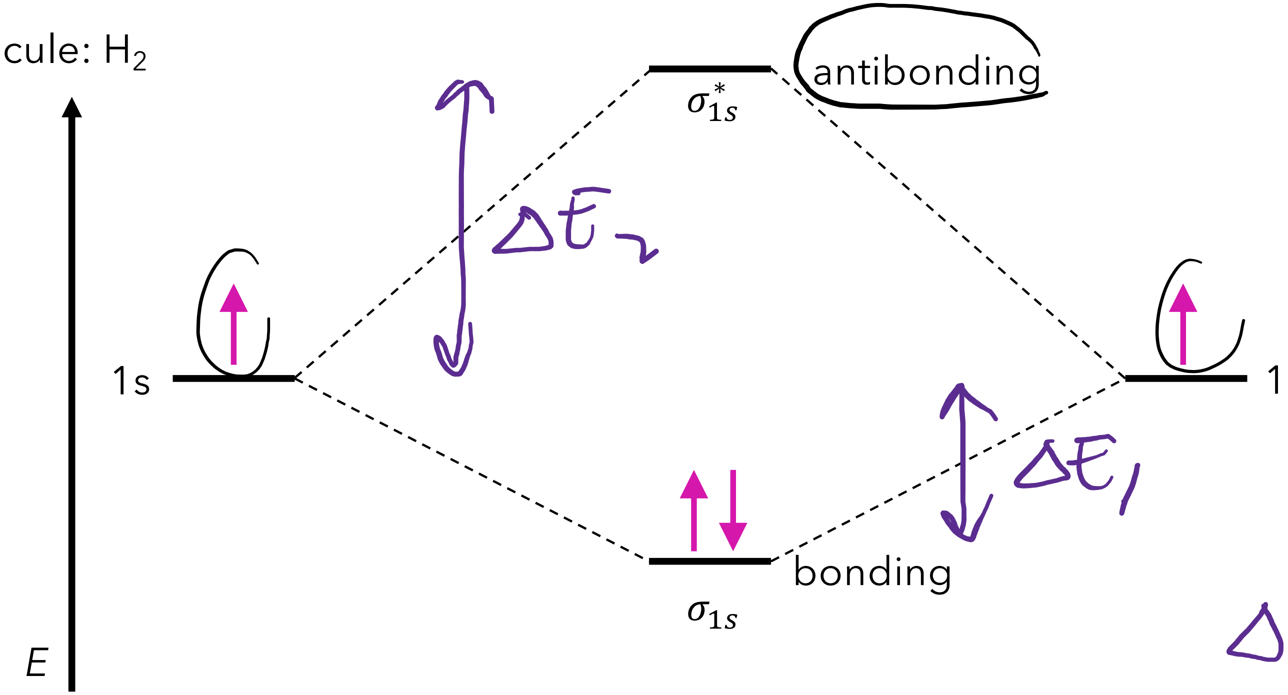
1. Form all the MOs that can be built from the available valence-shell atomic orbitals.

2. Populate them with the available electrons in the lowest-energy MO.

2G.1 Molecular orbitals

The energy of an antibonding orbital increased compared to the atomic orbitals!

Molecule: H_2



Electron configuration of H_2:
$(\sigma_{1s})^2$

↑ how many electrons

$\Delta E_2 > \Delta E_1$

The antibonding orbital is raised in energy more than the bonding orbital is lowered.
Two atomic 1s orbitals form **two molecular** σ_{1s} and σ^*_{1s} orbitals.

2G.1 Molecular orbitals

MO diagrams

- Increased energy of an **antibonding orbital is typically a little greater** than the lowering of the energy of the corresponding bonding orbital.
- Why? Although the bonding and antibonding orbitals have opposite effects on energy, the **repulsion between the nuclei** is the same in each case and pushes both orbital energies upward (Fig. 2G.4).

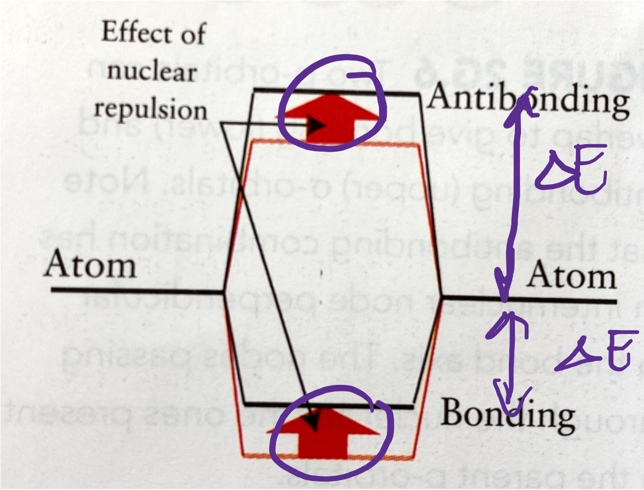
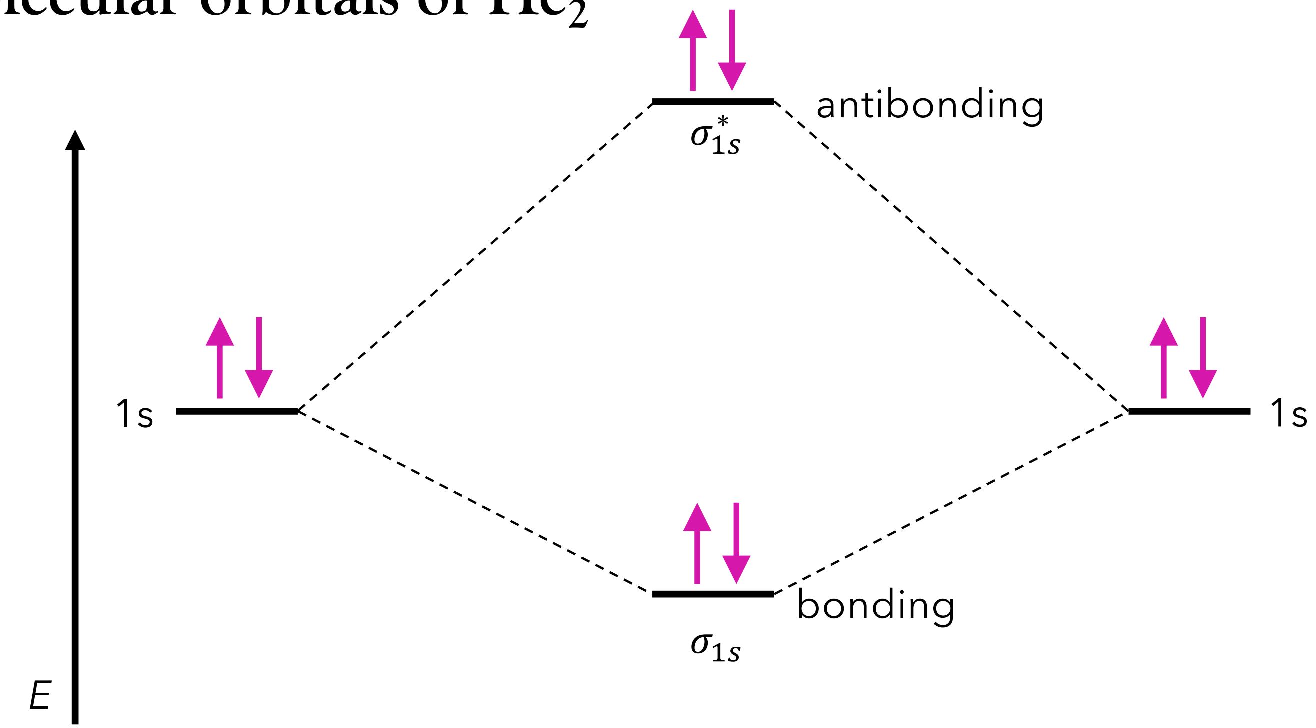


Figure 2G.4 (new book)

2G.1 Molecular orbitals

Molecular orbitals of He_2



Electron configuration of He_2:
$(\sigma_{1s})^2\ (\sigma_{1s}^*)^2$

Bond order (He_2) = ½(2-2) = 0
Bond order (H_2) = ½ (2-0) = 1

$$\boxed{\text{BOND ORDER} = b = \frac{1}{2} \times (N_e - N_e^*)}$$

Number of electrons in bonding orbitals: N_e
Number of electrons in antibonding orbitals: N_e^*

Important formula. Need to know.

2G.1 Molecular orbitals

Reality: He$_2$ does exist (barely)

He–He: weakest known bond. Discovered in 1993.

$$\Delta E_d = 0.01\frac{kJ}{mol} for\ He_2$$

$$\Delta E_d = 432\frac{kJ}{mol} for\ H_2$$

2G.2 Electron configurations of diatomic molecules

Procedure for determining the electronic configuration of diatomic molecules according to MO theory

Molecular orbitals are built from all the available valence-shell atomic orbitals (occupied or not). Valence electrons are then accommodated in these molecular orbitals following these rules:

1. Electrons occupy the **lowest-energy molecular orbital first**, then occupy orbitals of increasingly higher energy.
2. According to the Pauli exclusion principle, each molecular orbital can accommodate **up to two electrons**. If two electrons are present in one orbital, **their spins must be paired** (↑↓).
3. If more than one molecular orbital of the same energy is available, **the electrons enter them singly and adopt parallel spins** (Hund’s rule).

2G.2 Electron configurations of diatomic molecules

Valence-shell MOs for Period 1 & 2 homonuclear diatomic molecules

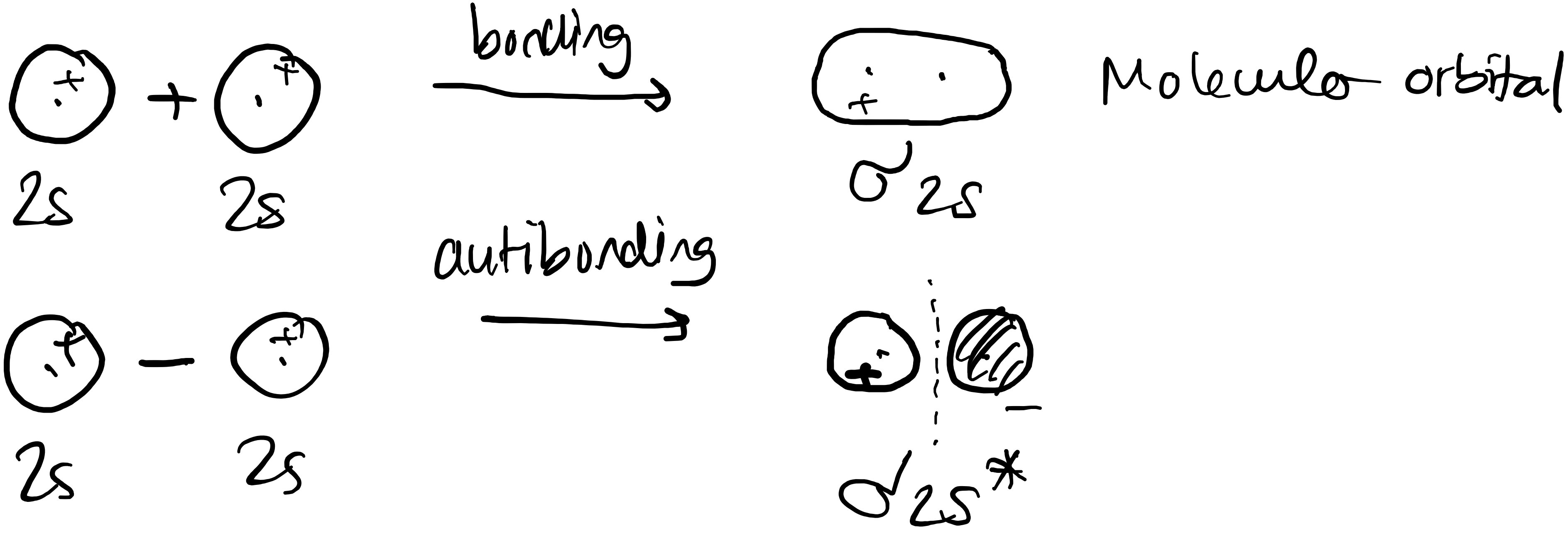
- Period 2: 2s and 2p orbitals available (four atomic orbitals)
- For diatomic molecules: total of eight AOs available to form eight MOs
- The 2s-orbitals overlap to form two σ-orbitals : one bonding (σ_{2s}) and one antibonding (σ_{2s}^{*}) (resemble σ_{1s} and σ_{1s}^{*} in hydrogen molecule).

2G.1 Molecular orbitals

Molecular Li_2 orbitals of

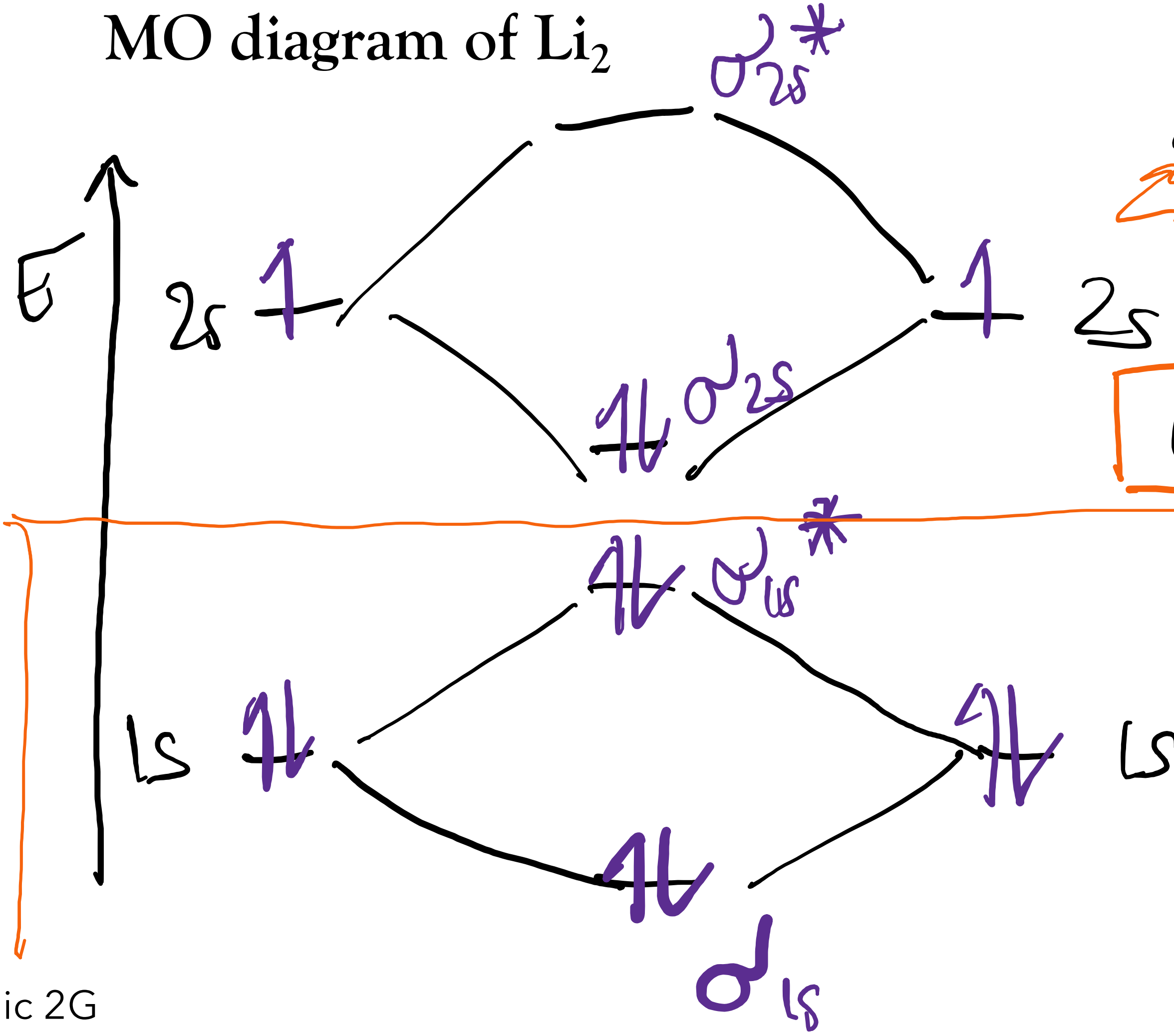
$Z = 3$, 1 valence

AO:



2G.1 Molecular orbitals

MO diagram of Li$_2$



$$BO(Li_2) = \frac{1}{2}(4-2) = 1$$

$$BO(Li_2) = \frac{1}{2}(2-0) = 1$$

← ignoring 1s

$$\boxed{(\sigma_{1s})^2(\sigma_{1s}^*)^2(\sigma_{2s})^2}$$

Addition:
In an exam,
We would tell you if you need to give the electron configuration of the entire molecule (incl. filled sigma-1s orbitals) or just the valence shell (without sigma-1s).

ignore this part!

→ I will not draw it again from here on out!

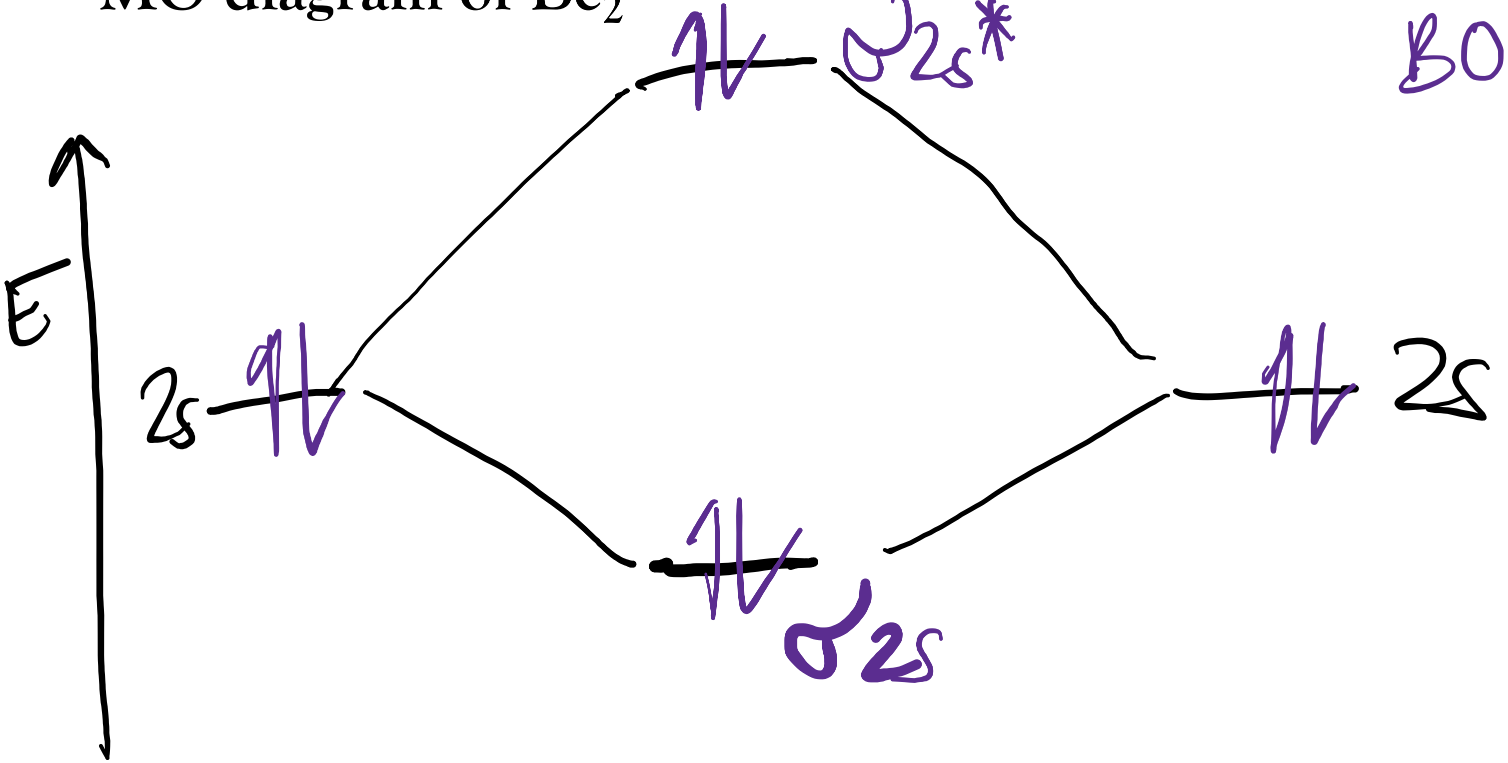
2G.1 Molecular orbitals

$Z = 4$ 2 valence e^-

MO diagram of Be_2

$BO = \frac{1}{2}(2-2) = 0$

$\Delta E_d = 9$ kJ/mol



2G.1 Molecular orbitals

Summary

Molecular orbitals are built from linear combinations of atomic orbitals: when atomic orbitals interfere **constructively**, they give rise to **bonding orbitals**; when they interfere **destructively**, they give rise to **antibonding orbitals**. N atomic orbitals combine to give N molecular orbitals.

2G.2 Electron configurations of diatomic molecules

Valence-shell MOs for Period 1 & 2 homonuclear diatomic molecules

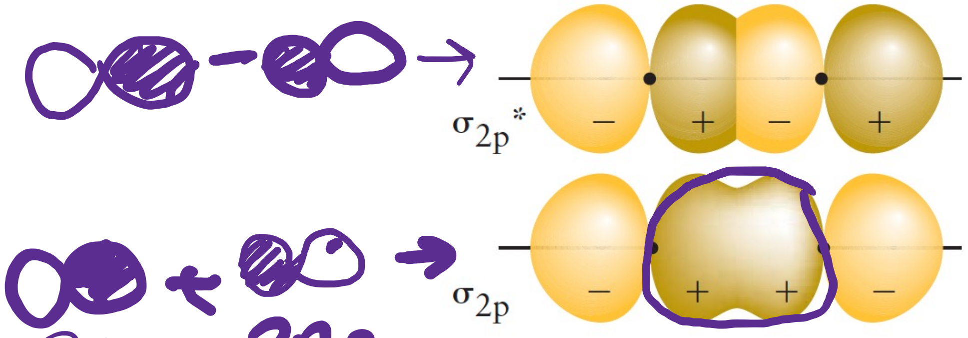
- Six 2p-orbitals (three on each atom)
- Overlap possible in two distinct ways:

1. End-to-end

Cylindrical symmetry, oriented along internuclear axis (one bonding, one antibonding)

2. Side-by-side:

- Oriented perpendicularly to internuclear axis
- Two bonding, two antibonding (only two shown here)



antibonding

bonding

2pz 2pz

Figure 3.29 (old book)

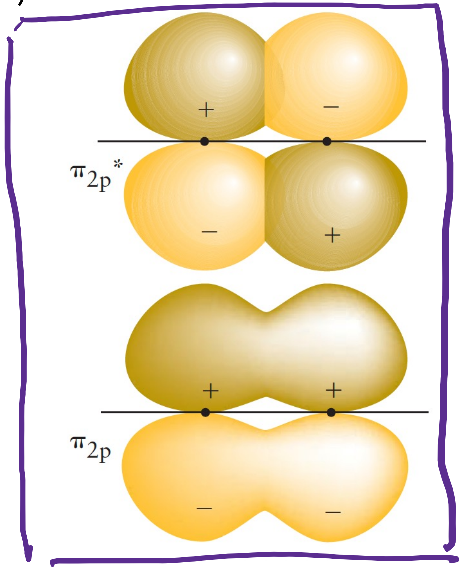
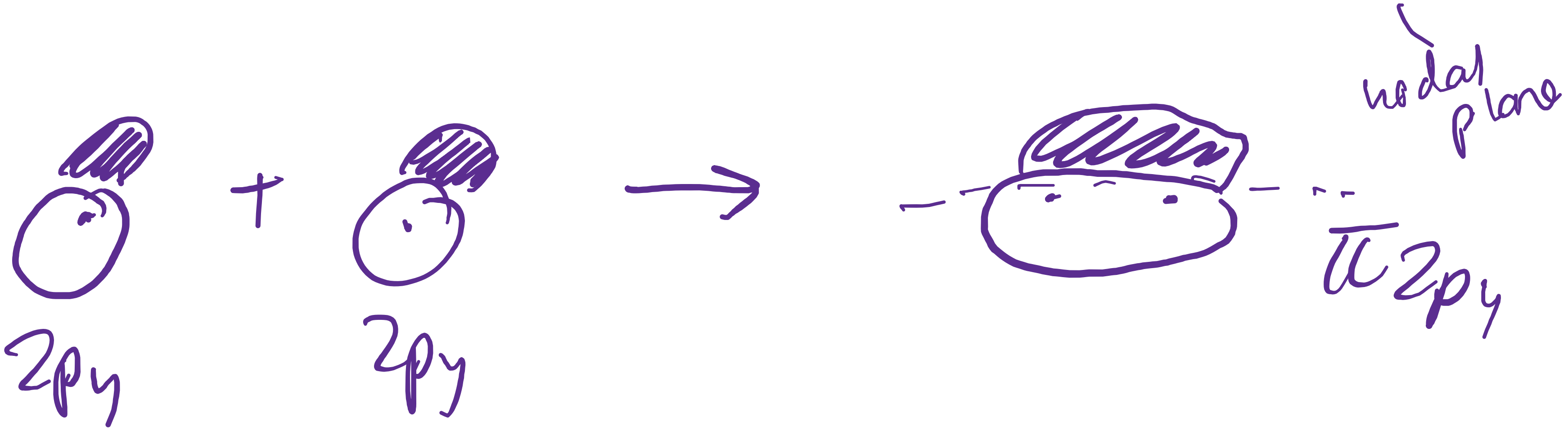
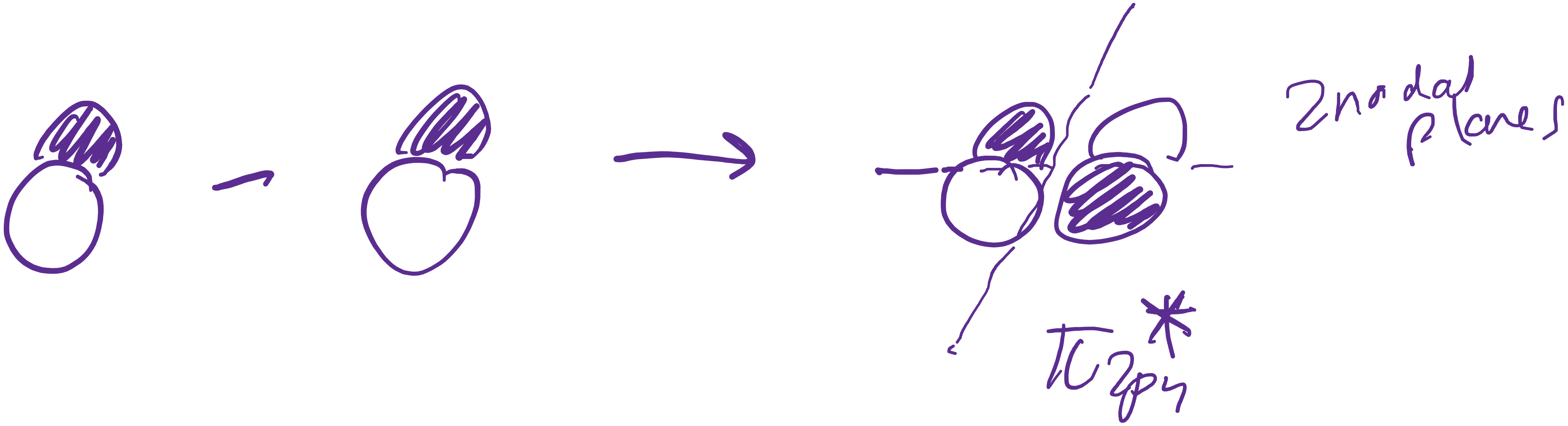


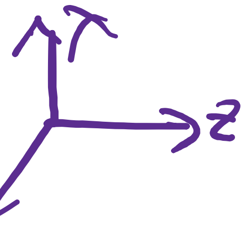
Figure 3.30 (old book)

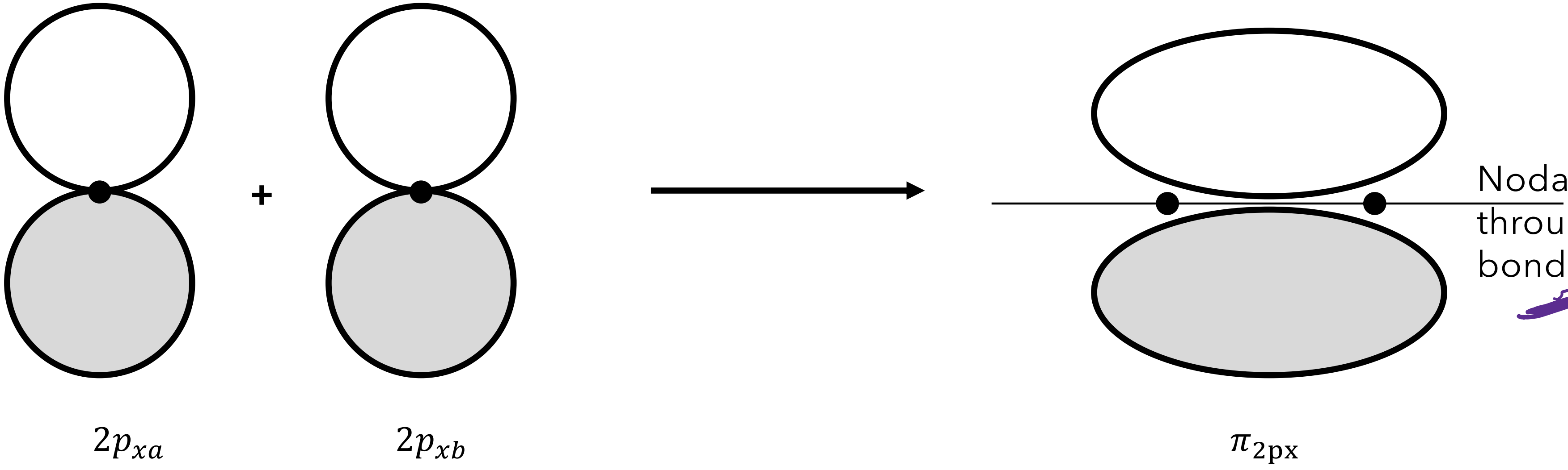

$2p_y$
+
$2p_y$
nodal plane
$\pi 2p_y$


−
2 nodal planes
$\pi^* 2p_y$

2G.2 Electron configuration of diatomic molecules

Mixing orbitals formed by LCAO of $2p_x$ and $2p_y$ via constructive interference



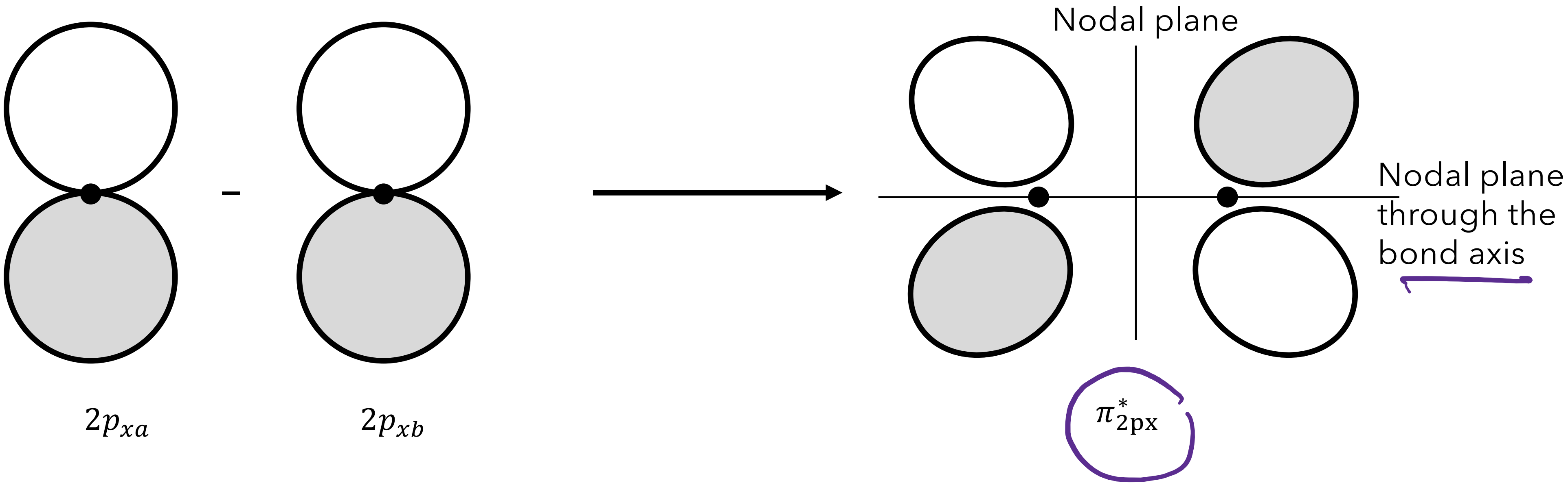


Nodal plane through the bond axis

π-orbital: Molecular orbital with a nodal plane through the bond axis.

2G.2 Electron configuration of diatomic molecules

Mixing orbitals formed by LCAO of 2p$_x$ and 2p$_y$ via destructive interference



π^*-orbital: Molecular orbital with two nodal planes.

2G.2 Electron configurations of diatomic molecules

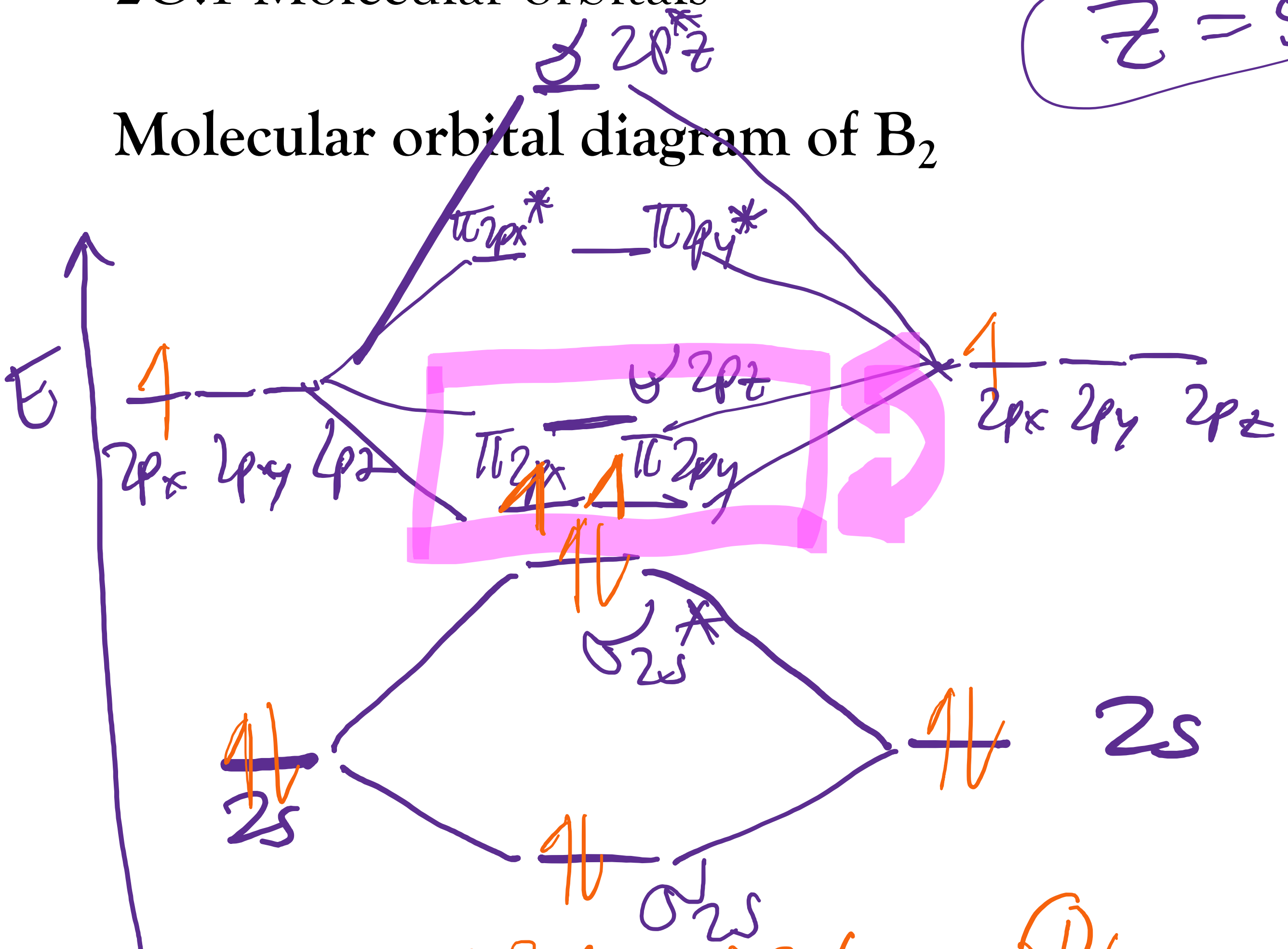
σ- and π-orbitals

- A **σ-orbital** is a molecular orbital that has **cylindrical symmetry** and **no nodal plane that contains the internuclear axis**.
- A **π-orbital** is a molecular orbital with a **nodal plane that contains the internuclear axis**.
- Both types of orbitals can be either bonding or antibonding:

  Bonding orbitals do not have internuclear nodes arising from destructive interference, antibonding orbitals do.

2G.1 Molecular orbitals

Molecular orbital diagram of B_2



$\boxed{Z = 5}$, 3 valence e^-

$Z \geq 8 \rightarrow \sigma, \pi$

$Z < 8 \rightarrow \pi, \sigma$

2 unpaired e^-

paramagnetic!

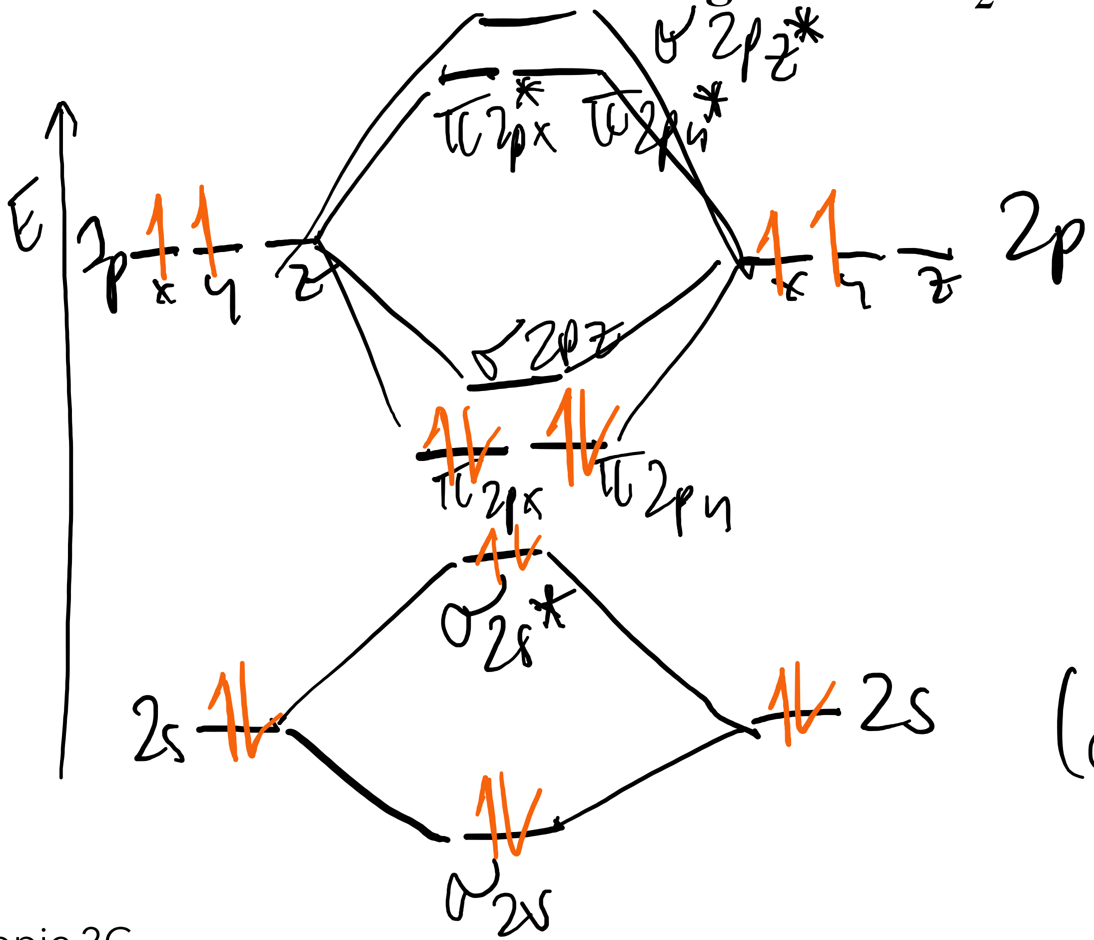
$BO = \frac{1}{2}(4-2) = 1$

single bond

$(\sigma_{2s})^2 (\sigma^*_{2s})^2 (\pi 2p_x)^{①} (\pi 2p_y)^1$

2G.1 Molecular orbitals

Molecular orbital diagram of C_2



$Z = 6 \rightarrow \pi$ first, then σ

C has 4 $e^{\ominus}$

Bond order $= \frac{1}{2}(6-2) = 2$

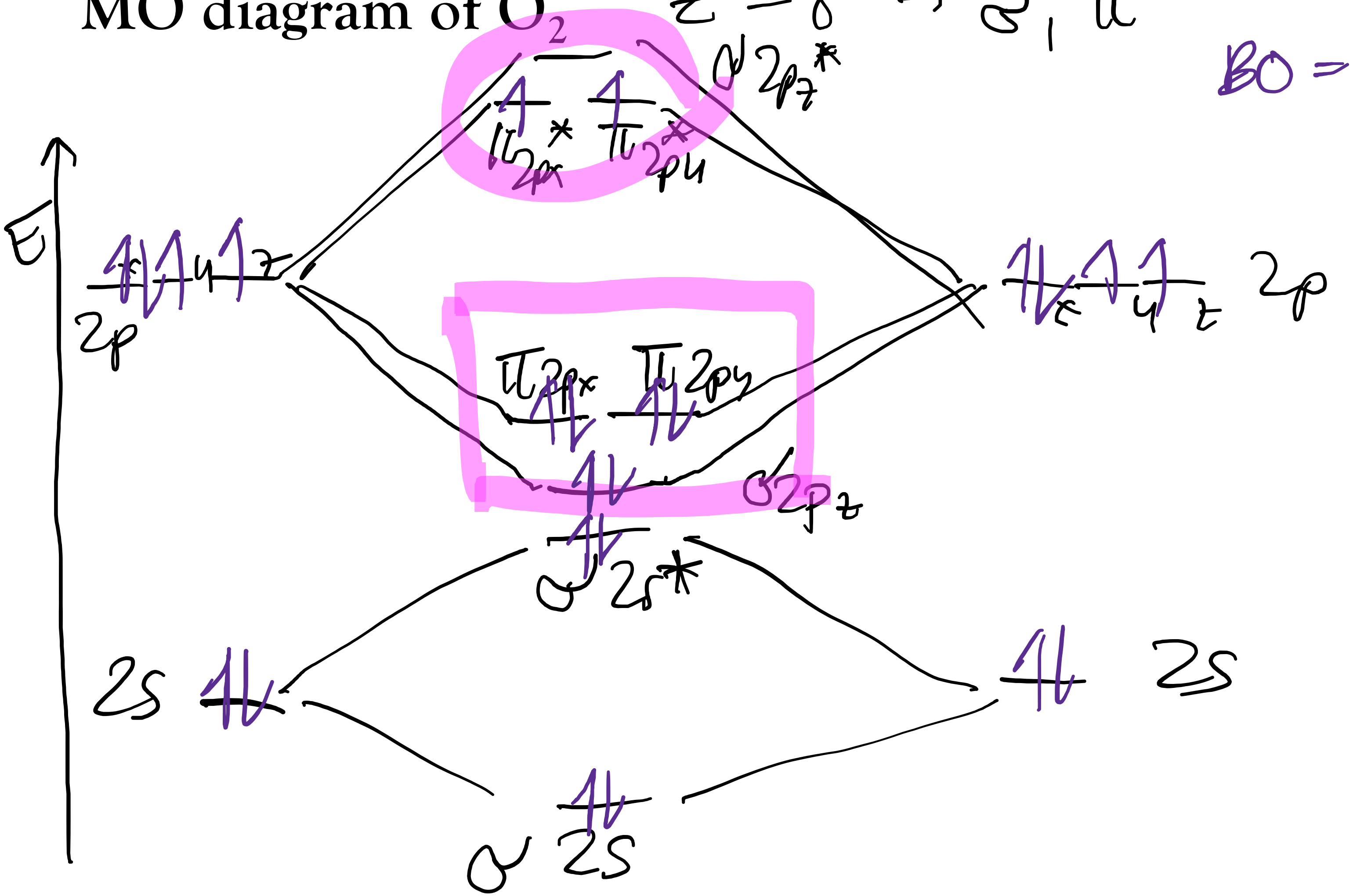
double bond

Valence shell electron configuration:

$$(\sigma_{2s})^2(\sigma_{2s}^*)^2\underbrace{(\pi_{2px})^2(\pi_{2py})^2}_{=(\pi_{2p})^4}$$

MO diagram of O_2

$Z = 8 \rightarrow \sigma, \pi$



$BO = \frac{1}{2}(8 - 4) = \underline{\underline{2}}$

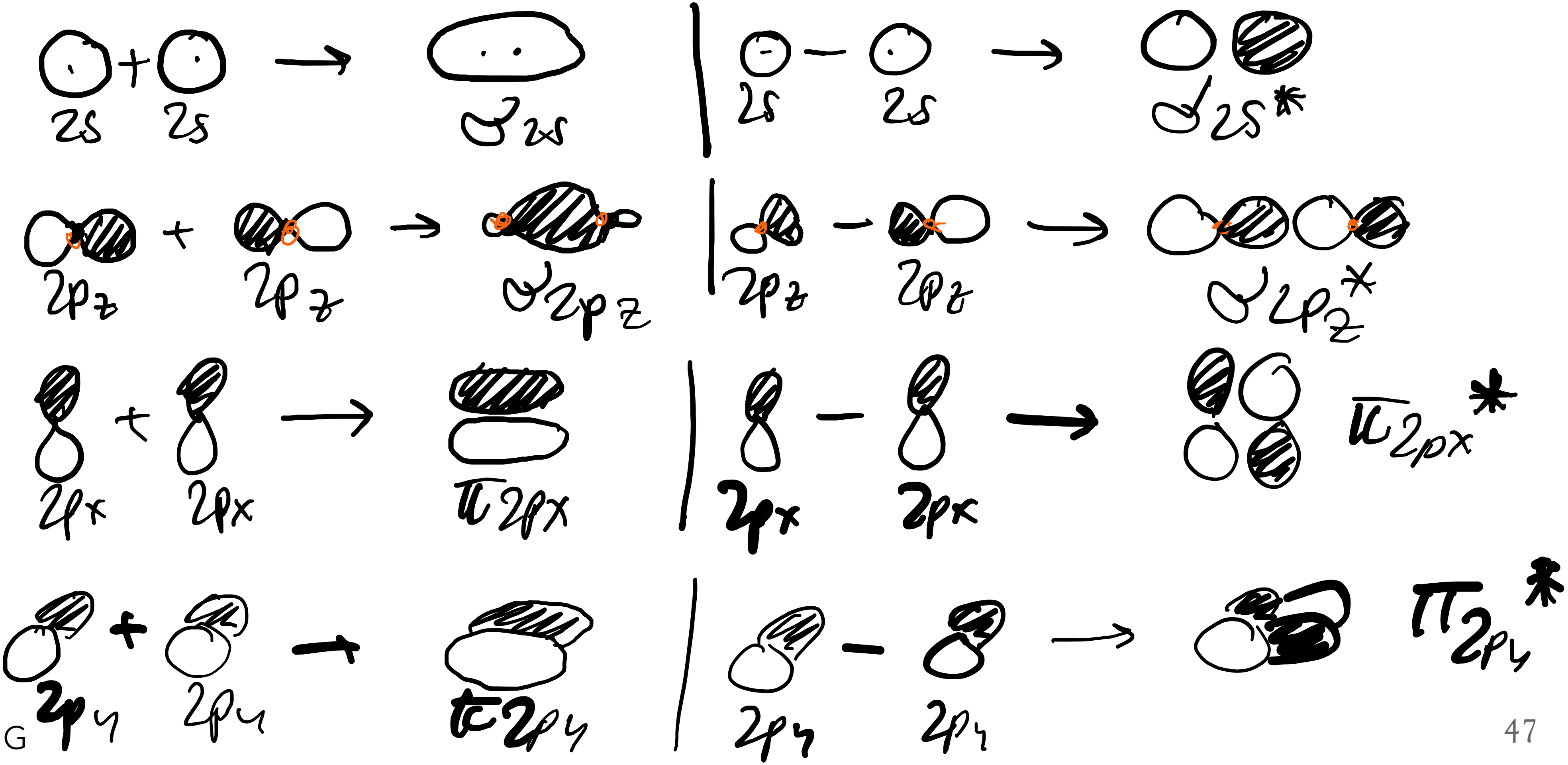
double bond

2G.1 Molecular orbitals

$Z=8 \rightarrow \sigma, \pi$, 6 valence e^-

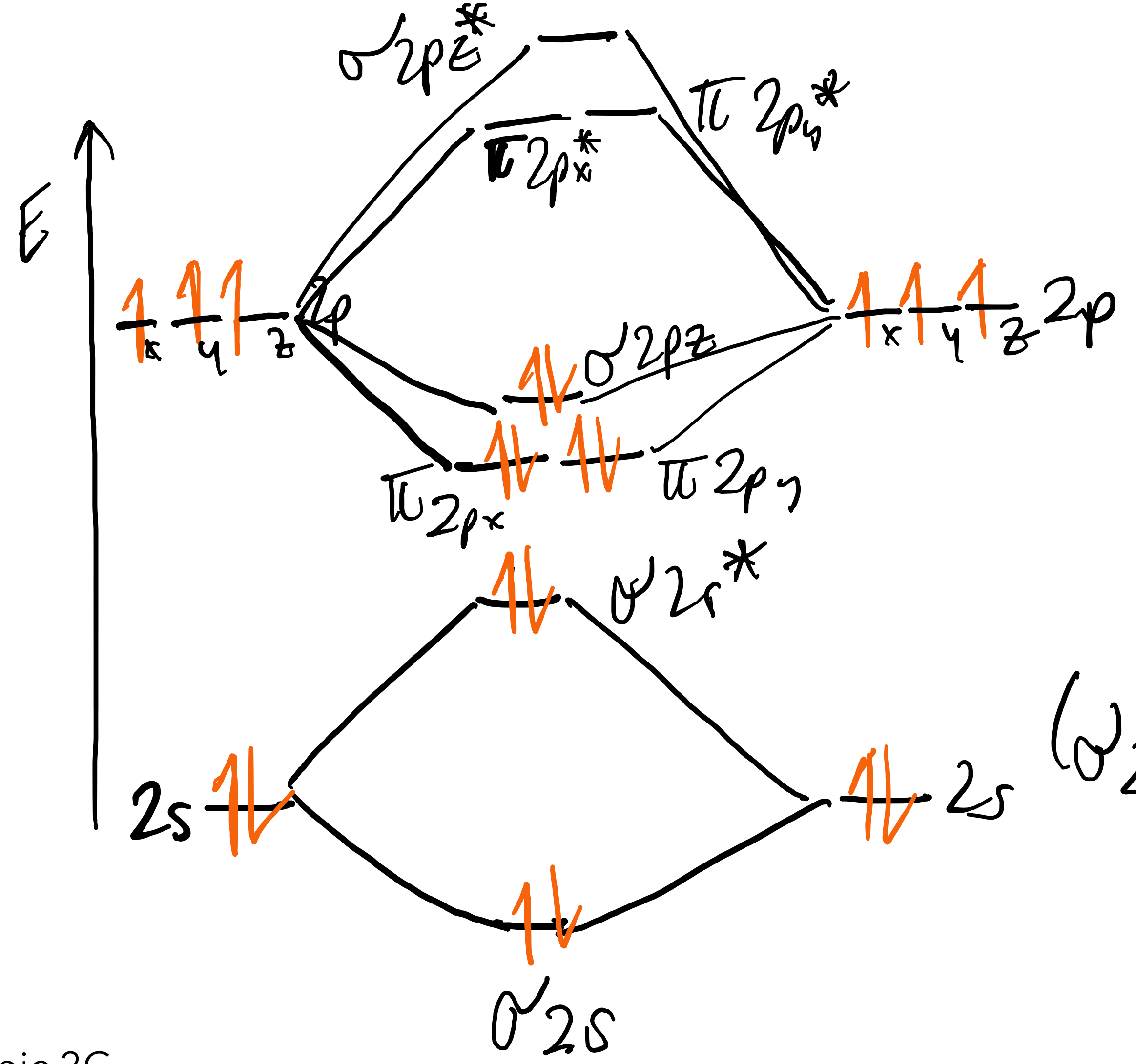
AO mix: 2s, 3 × 2p

Molecular orbitals of O_2



MO diagram of N_2

Z=7 → π first. N has 5 valence e^{-}



$BO = \frac{1}{2}(8-2) = 3$

triple bond

Electron config.
(valence shell!):
(no 1s)

$(\sigma_{2s})^2(\sigma_{2s}^*)^2(\pi_{2p_x})^2(\pi_{2p_y})^2(\sigma_{2p_z})^2$

2G.2 Electron configurations of diatomic molecules

MO diagrams for homonuclear diatomic molecules

Li_2 to N_2

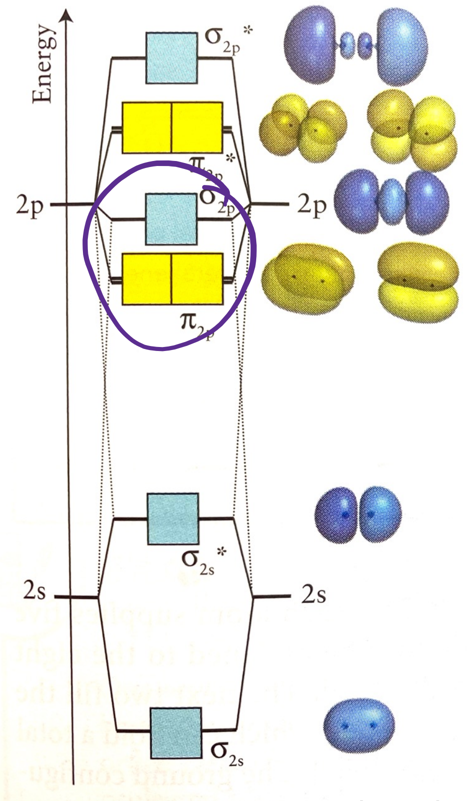


Figure 2G.8 (new book)

O_2 and F_2

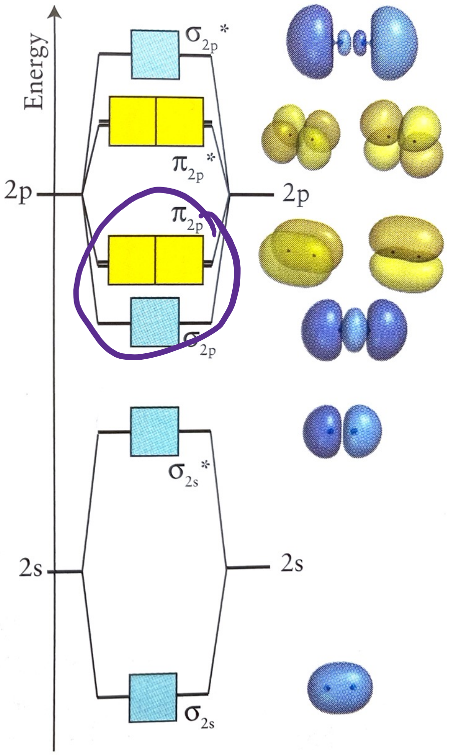
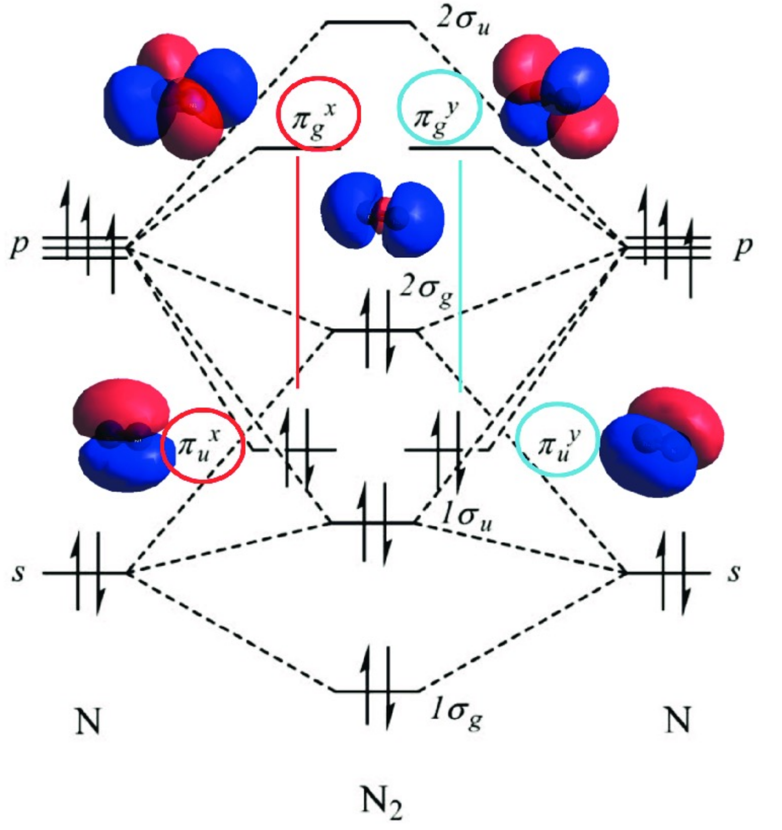


Figure 2G.9 (new book)

- Note: order of mixed molecular orbitals is different!
- Up to Z=7, pie first ☺
- At mature age of Z=8, pie second.

2G.2 Electron configurations of diatomic molecules

MO diagrams for homonuclear diatomic molecules



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2G.2 Electron configurations of diatomic molecules

MO diagrams for homonuclear diatomic molecules

O_2 and F_2

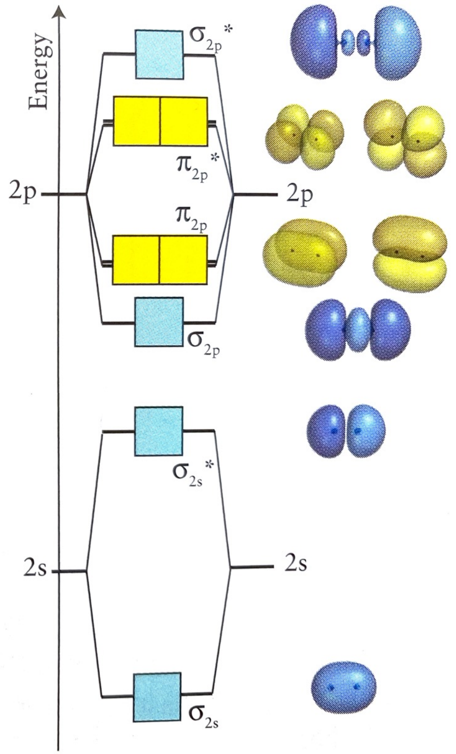


Figure 2G.9 (new book)

- Note: order of mixed molecular orbitals is different!
- **O and F atoms** have many electrons that contribute to shielding. 2s and 2p orbitals are well separated.
- Order as shown:

$$\sigma_{2s} < \sigma_{2s}^* < \sigma_{2p} < \pi_{2p} < \pi_{2p}^* < \sigma_{2p}^*$$

2G.2 Electron configurations of diatomic molecules

MO diagrams for homonuclear diatomic molecules

Li_2 to N_2

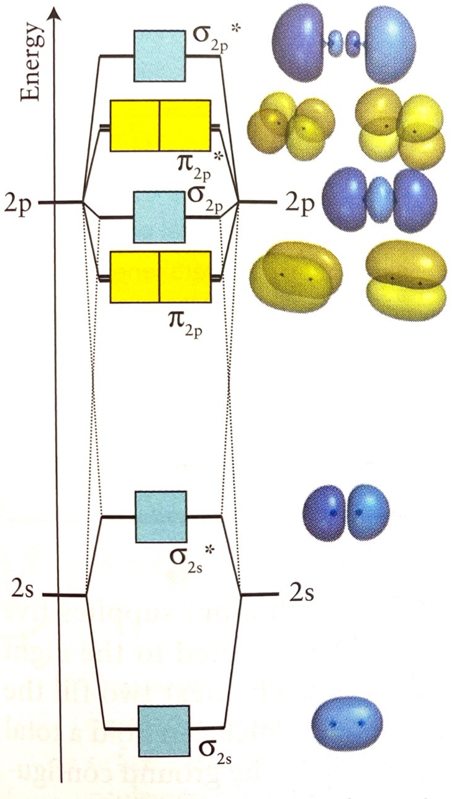


Figure 2G.8 (new book)

- Note: order of mixed molecular orbitals is different!
- **For Li to N**, these atoms have fewer electrons, less shielding: the 2s and 2p orbitals are similar in energy.
- The σ-orbitals for these molecules are formed by mixing both 2s/2p orbitals. **It is difficult to predict where the resulting σ-orbitals lie. <u>Need to know by heart.</u>**
- Experimentally, they turn out to lie as shown in Figure 2.8:

$$\sigma_{2s} < \sigma_{2s}^{*} < \pi_{2p} < \sigma_{2p} < \pi_{2p}^{*} < \sigma_{2p}^{*}$$

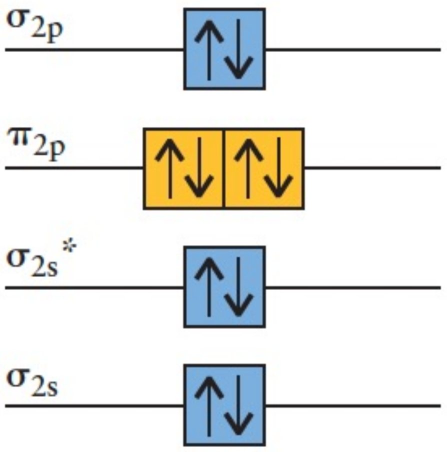
2G.2 Electron configurations of diatomic molecules

Ground-state electron configurations in MO theory

- Use building-up principle to accommodate all the valence electrons
- Each atom supplies five valence electrons
- Ten electrons assigned to orbitals shown in Figure 2G.8
- Electron configuration in ground state:

$$N_2: (\sigma_{2s})^2(\sigma^*_{2s})^2\underbrace{(\pi_{2p})^4}_{=(\pi_{2px})^2(\pi_{2py})^2}(\sigma_{2p})^2$$

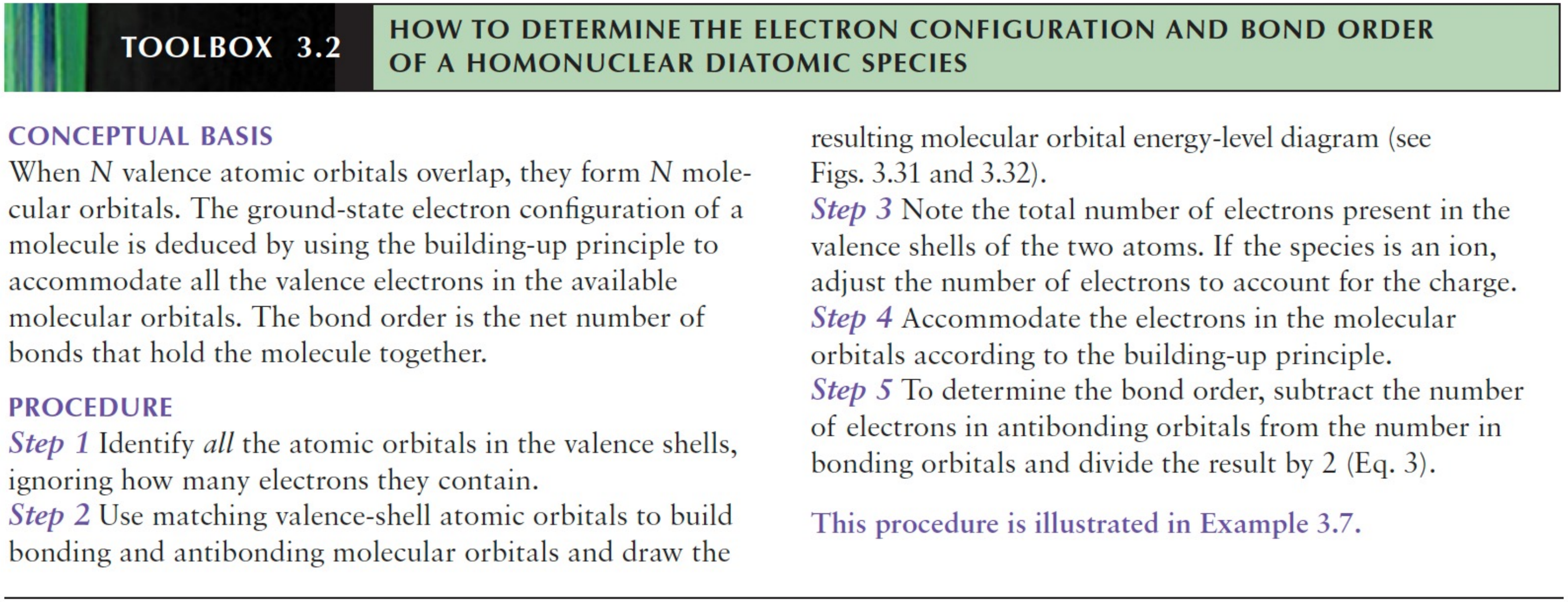
both okay!



45 Nitrogen, N_2

2G.2 Electron configurations of diatomic molecules

Electron configuration and bond order in MO theory



TOOLBOX 3.2 **HOW TO DETERMINE THE ELECTRON CONFIGURATION AND BOND ORDER OF A HOMONUCLEAR DIATOMIC SPECIES**

CONCEPTUAL BASIS

When *N* valence atomic orbitals overlap, they form *N* molecular orbitals. The ground-state electron configuration of a molecule is deduced by using the building-up principle to accommodate all the valence electrons in the available molecular orbitals. The bond order is the net number of bonds that hold the molecule together.

PROCEDURE

Step 1 Identify *all* the atomic orbitals in the valence shells, ignoring how many electrons they contain.

Step 2 Use matching valence-shell atomic orbitals to build bonding and antibonding molecular orbitals and draw the resulting molecular orbital energy-level diagram (see Figs. 3.31 and 3.32).

Step 3 Note the total number of electrons present in the valence shells of the two atoms. If the species is an ion, adjust the number of electrons to account for the charge.

Step 4 Accommodate the electrons in the molecular orbitals according to the building-up principle.

Step 5 To determine the bond order, subtract the number of electrons in antibonding orbitals from the number in bonding orbitals and divide the result by 2 (Eq. 3).

This procedure is illustrated in Example 3.7.

2G.2 Electron configurations of diatomic molecules

Summary

The ground-state electron configurations of diatomic molecules are deduced by forming molecular orbitals from all the valence-shell atomic orbitals of the two atoms and adding the valence electrons to the molecular orbitals in order of increasing energy, in accord with the building-up principle.

2G.3 Bonding in heteronuclear diatomic molecules

more electronegative
↓
lower in energy

Heteronuclear diatomic molecules

- Use same approach as for homonuclear diatomic molecules
- **Energy of two atoms will be different**

X=C

- Mix 2s and $2p_z$ orbitals of both atoms to create σ-orbitals
- Mix $2p_x$ and $2p_y$ orbitals of both atoms to create π-orbitals
- Relative energies are hard to estimate (**you don't need to know!**), but can be calculated
- The relative energies of the resulting MOs for NO and CO (X= N or C) are shown in Figure 2G.12

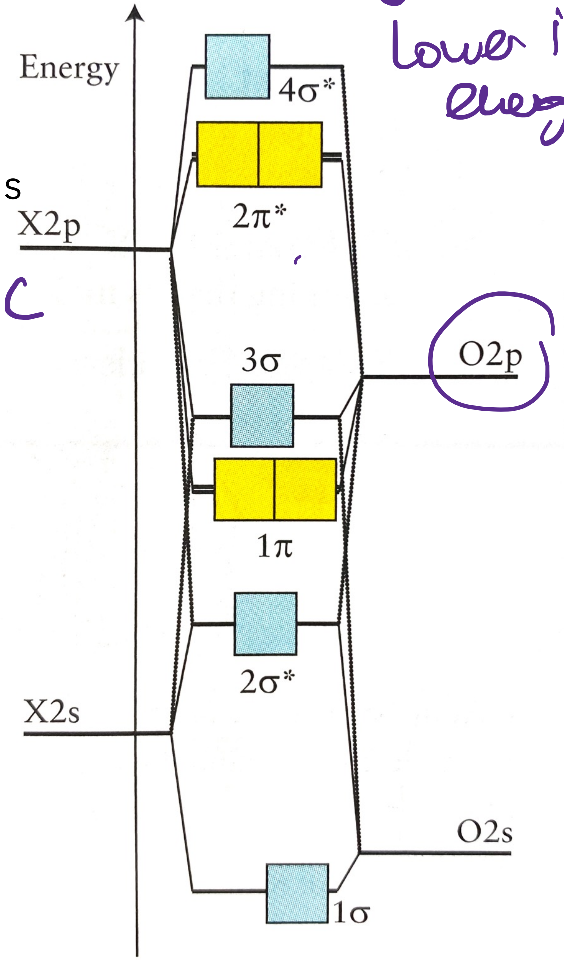


Figure 2G.12 (new book)

MO diagram of CO

Will not be part of exam!

2G.3 Bonding in heteronuclear diatomic molecules

Summary

Bonding in heteronuclear diatomic molecules involves an unequal sharing of the bonding electrons. The more electronegative element contributes more strongly to the bonding orbitals, whereas the less electronegative element contributes more strongly to the antibonding orbitals.

2G.5 A comparison of bonding models

Table 2G.1

homonuclear diatomic molecules

↑

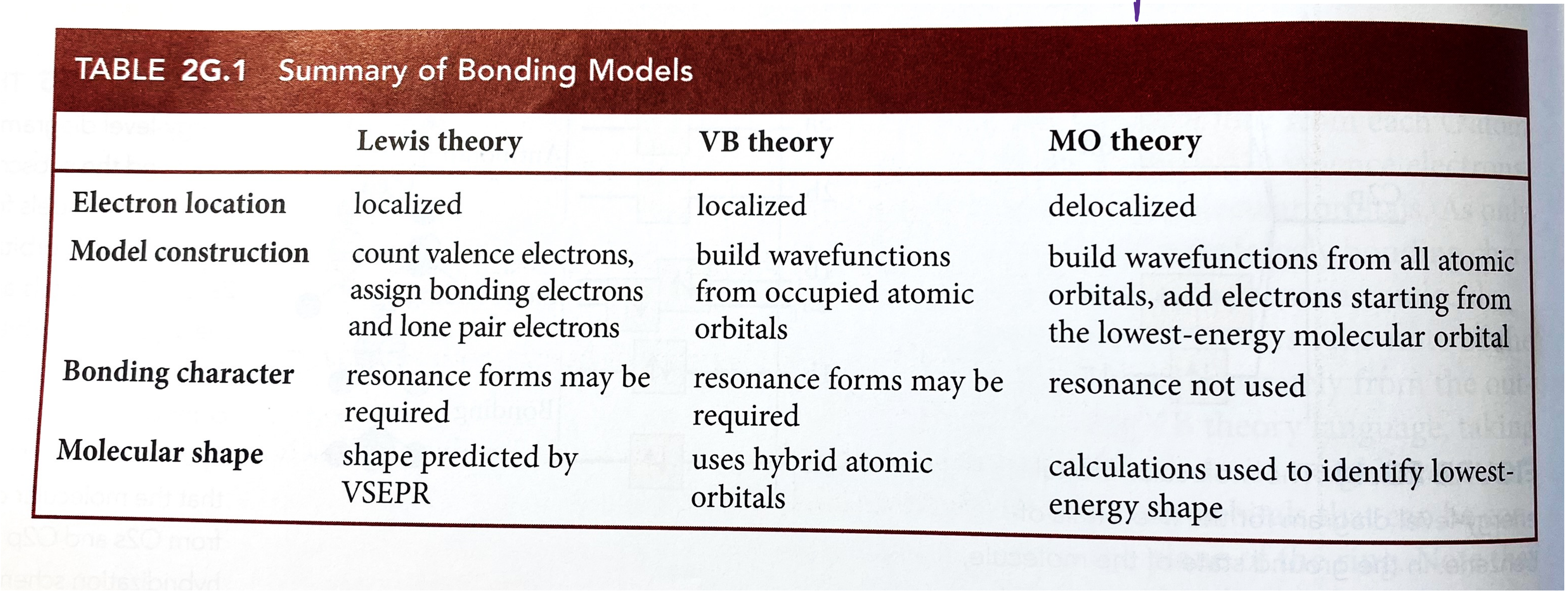


TABLE 2G.1 Summary of Bonding Models

	Lewis theory	VB theory	MO theory
Electron location	localized	localized	delocalized
Model construction	count valence electrons, assign bonding electrons and lone pair electrons	build wavefunctions from occupied atomic orbitals	build wavefunctions from all atomic orbitals, add electrons starting from the lowest-energy molecular orbital
Bonding character	resonance forms may be required	resonance forms may be required	resonance not used
Molecular shape	shape predicted by VSEPR	uses hybrid atomic orbitals	calculations used to identify lowest-energy shape

The skills you have mastered are the ability to

- Construct and interpret a molecular orbital energy-level diagram for homonuclear diatomic molecules
- Deduce the ground-state electron configurations of Period 2 diatomic molecules
- Define and use bond order as an assessment of the number of bonds between pairs of atoms

Summary: You have learned that according to MO theory, bonding is described by wavefunctions (molecular orbitals) that spread over all the atoms in a molecule and that each orbital can be occupied by up to two electrons. You now know about the existence of σ- and π-orbitals and bonding, antibonding, and nonbonding orbitals. Their systematic occupation according to the building-up principle is used to predict the ground-state electron configuration of a molecule. You have seen how MO theory accounts for the paramagnetism of some molecules.