

# CHAPTER 14: COVALENT BONDING: ORBITALS

## PROBLEMS WITH LEWIS STRUCTURE

- Provides no indication of the **shapes** of the molecules
- $O_2$  would be predicted to be diamagnetic, but it is **paramagnetic**

**Paramagnetic** — has unpaired electrons.



The Lewis structure of  $O_2$  has an  $O=O$  double bond with no unpaired electrons, i.e. diamagnetic.

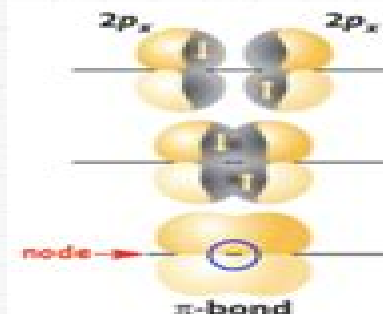
- Many compounds cannot be represented by a single structure
- Resonance structures **differ only in position of electrons** not the nuclei.
- Result in notions of hypervalent compounds and "expansion of octet"
- Can't explain molecules such as  $B_2H_6$  (electron deficient molecules)

## VALENCE-BOND THEORY "LOCALIZED ELECTRON MODEL"

- Covalent bond is created by the overlap of orbitals on two atoms

### TYPES OF OVERLAP: $\sigma$ and $\pi$ Bonds

- **Bonding in  $H_2$** 
  - Sausage shaped orbital, bond called  **$\sigma$ -orbital**
  - Called  $\sigma$  orbital because it has **no node and is cylindrically symmetrical**
  - Electron density is **concentrated between two nuclei**, but is still highest at the nuclei themselves.
- **Bonding in  $N_2$** 
  - **$\sigma$  orbital** is obtained by overlap of **two  $p_z$ -orbitals**
- Perpendicular p orbitals overlap in a side-by-side manner.
- The resulting orbital has two lobes of electron density and is called a  **$\pi$ -bond**.
- Called  $\pi$ -orbital because it **has a node** and, when viewed down the internuclear axis, it looks like a p-orbital
- $\pi$ -bond has **no electron density between the nuclei along the bond axis**.



# General Chemistry Chapter 14 Covalent Bonding Orbitals Exercises

**Kenneth W. Whitten**



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