

CHAPTER 14: COVALENT BONDING: ORBITALS

PROBLEMS WITH LEWIS STRUCTURE

- Provides no indication of the **shapes** of the molecules
- O₂ would be predicted to be diamagnetic, but it is **paramagnetic**

Paramagnetic = has unpaired electrons.



The Lewis structure of O₂ has an O=O double bond with **no unpaired electrons**, i.e. diamagnetic.

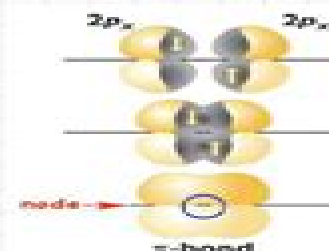
- Many compounds cannot be represented by a single structure
- Resonance structures **differ only in position of electrons** not the nuclei.
- Result in notions of hypervalent compounds and "expansion of octet"
- Can't explain molecules such as B₂H₆ (electron deficient molecules)

VALENCE-BOND THEORY "LOCALIZED ELECTRON MODEL"

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

TYPES OF OVERLAP: σ and π Bonds

- Bonding in H₂**
 - Sausage shaped orbital, bond called **σ -orbital**
 - Called σ 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 N₂
 - σ orbital** is obtained by overlap of **two p_x-orbitals**
- Perpendicular p orbitals overlap in a side-by-side manner
- The resulting orbital has two lobes of electron density and is called a **π -bond**
- Called π -orbital because it **has a node** and, when viewed down the internuclear axis, it looks like a p-orbital
- π -bond has **no electron density between the nuclei along the bond axis**



General Chemistry Chapter 14 Covalent Bonding Orbitals Practice

Kaplan Test Prep

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