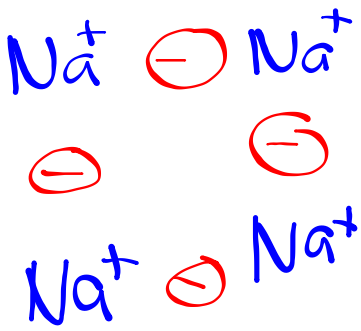
Unit 3 - Chemical Bonding

- Electron Configurations
- Octet Rule
- Electron Dot Structure
- Metallic Bonding
- Covalent Bonding
- VSEPR Theory
- Hybridization
- Polarity
- Intermolecular Forces


Na+
Na+
Na+
Na+

Quantum Mechanical Model of an Atom

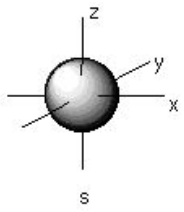
The quantum mechanical model determines the allowed energies an electron can have and how likely it is to find the electron in various locations around the nucleus.

atomic orbital - region of space in which there is a high probability to find an electron

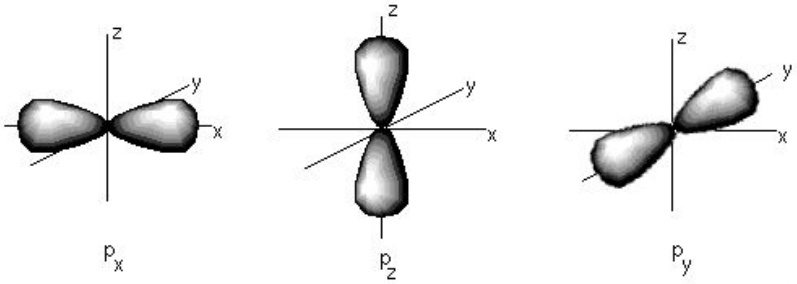
Principal quantum numbers (**n**) represent energy levels of electrons (i.e., n = 1, 2, 3, 4, etc.)

There may be several orbitals with different shapes at different energy levels.

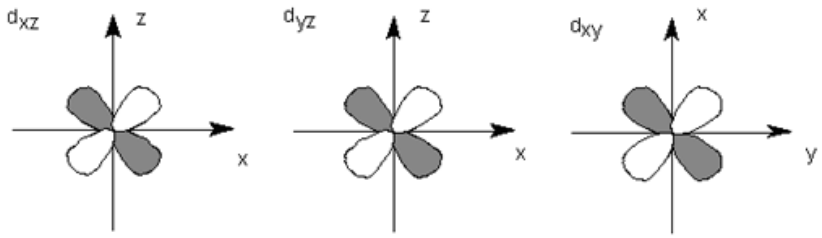
s orbital

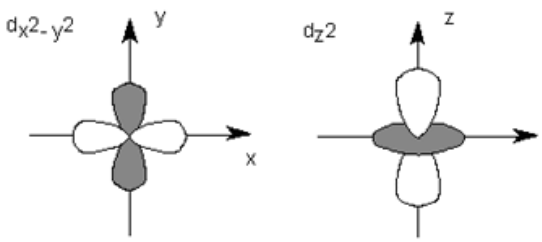


p orbitals

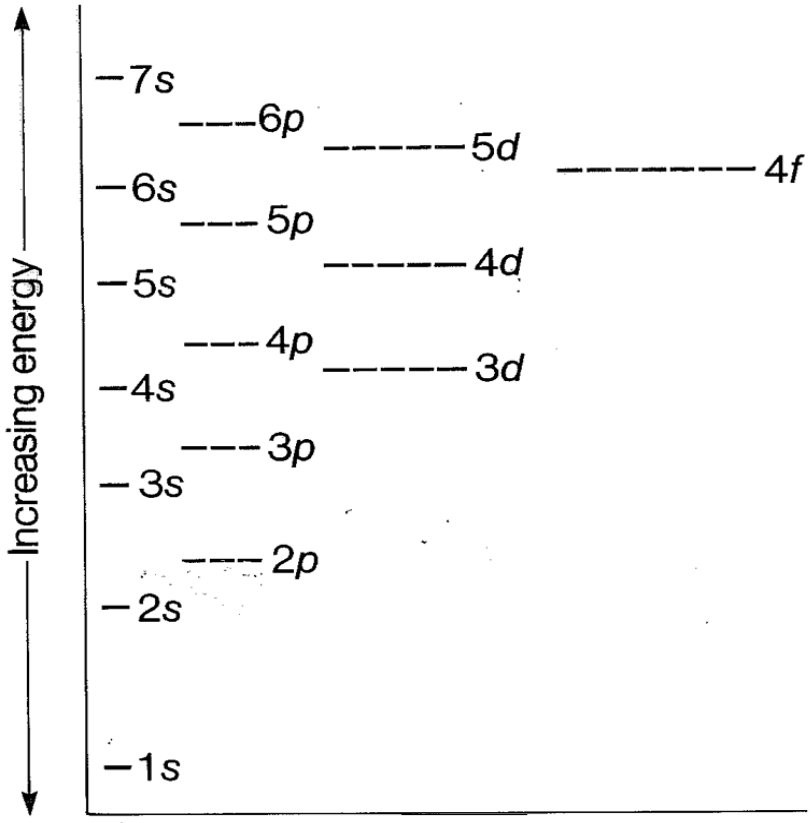


d orbitals





Aufbau Diagram



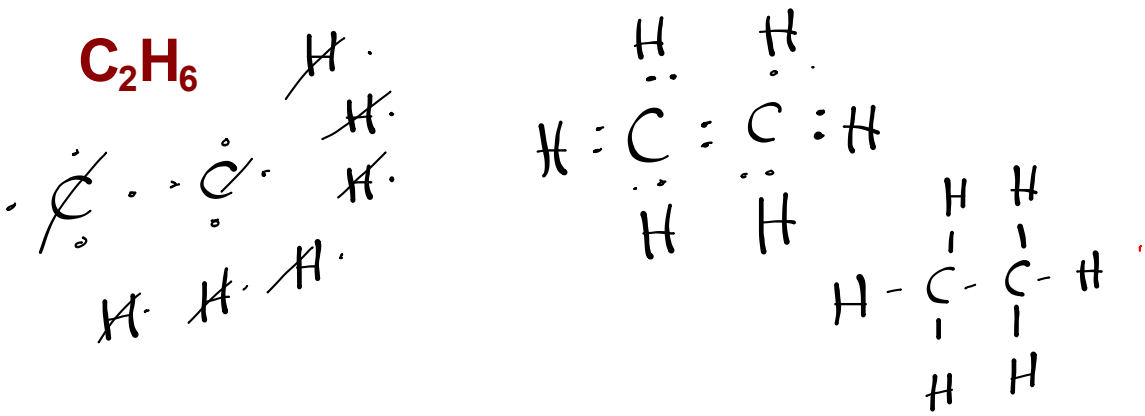
Aufbau principle - electrons occupy orbitals of lowest energy first

Pauli exclusion principle- an atomic orbital can describe at most two electrons

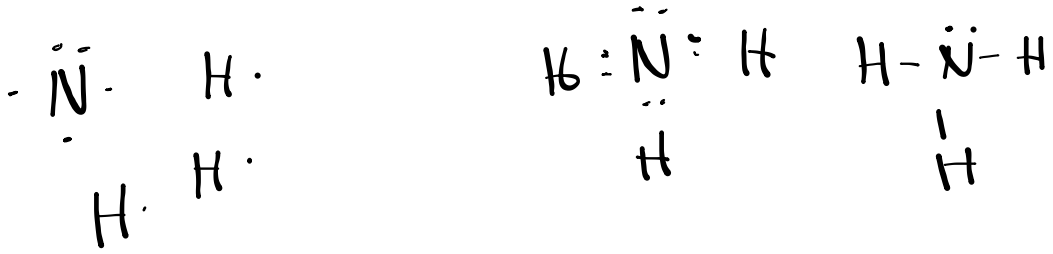
Hund's rule - one electron enters each orbital until all orbitals contain one electron with the same spin

Electron Dot Structures and Hybridization

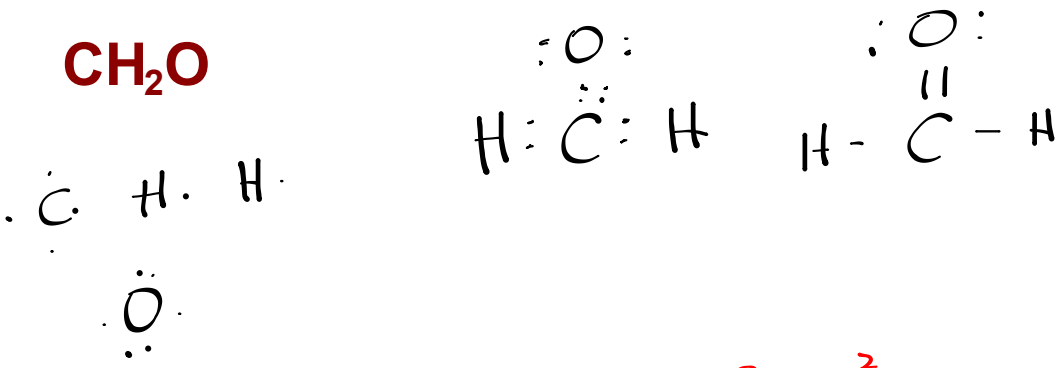
C_2H_6



NH_3



CH_2O



$S + p_x + p_y \rightarrow sp^2 + sp^2 + sp^2$

σ

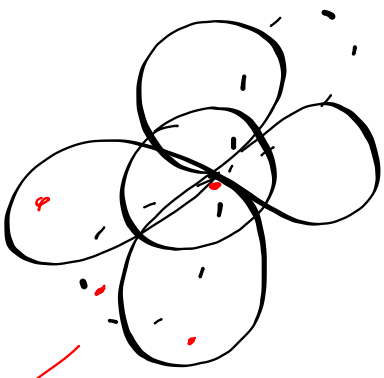
$p_z \rightarrow p_z$

π bond

CH_4 ← H· H· H· H·

2p 1 1 1 ·C·

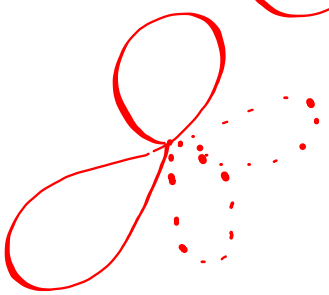
2s 1

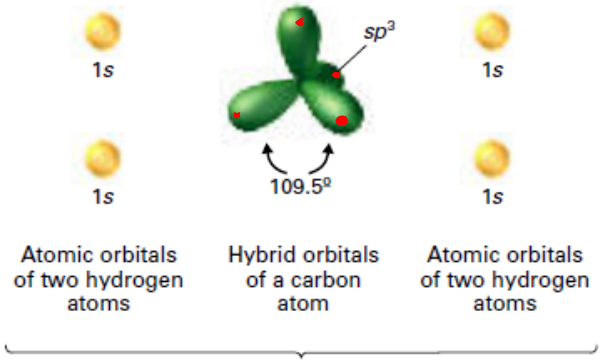


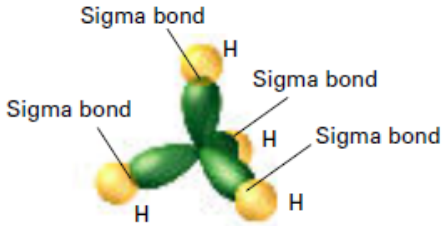
hybridization

$S + p_x + p_y + p_z \rightarrow sp^3 + sp^3 + sp^3 + sp^3$

σ bonds







Methane molecule

Table 8.3 Electronegativity Differences and Bond Types

Electronegativty difference range	Most probable type of bond	Example
0.0-0.4	Nonpolar covalent	H - H (0.0)
0.4-1.0	Moderately polar covalent	H - Cl (0.9)
1.0-2.0	Very polar covalent	H - F (1.9)
≥ 2.0	Ionic	Na^+ Cl^- (2.1)

* No sharp boundary between ionic and covalent

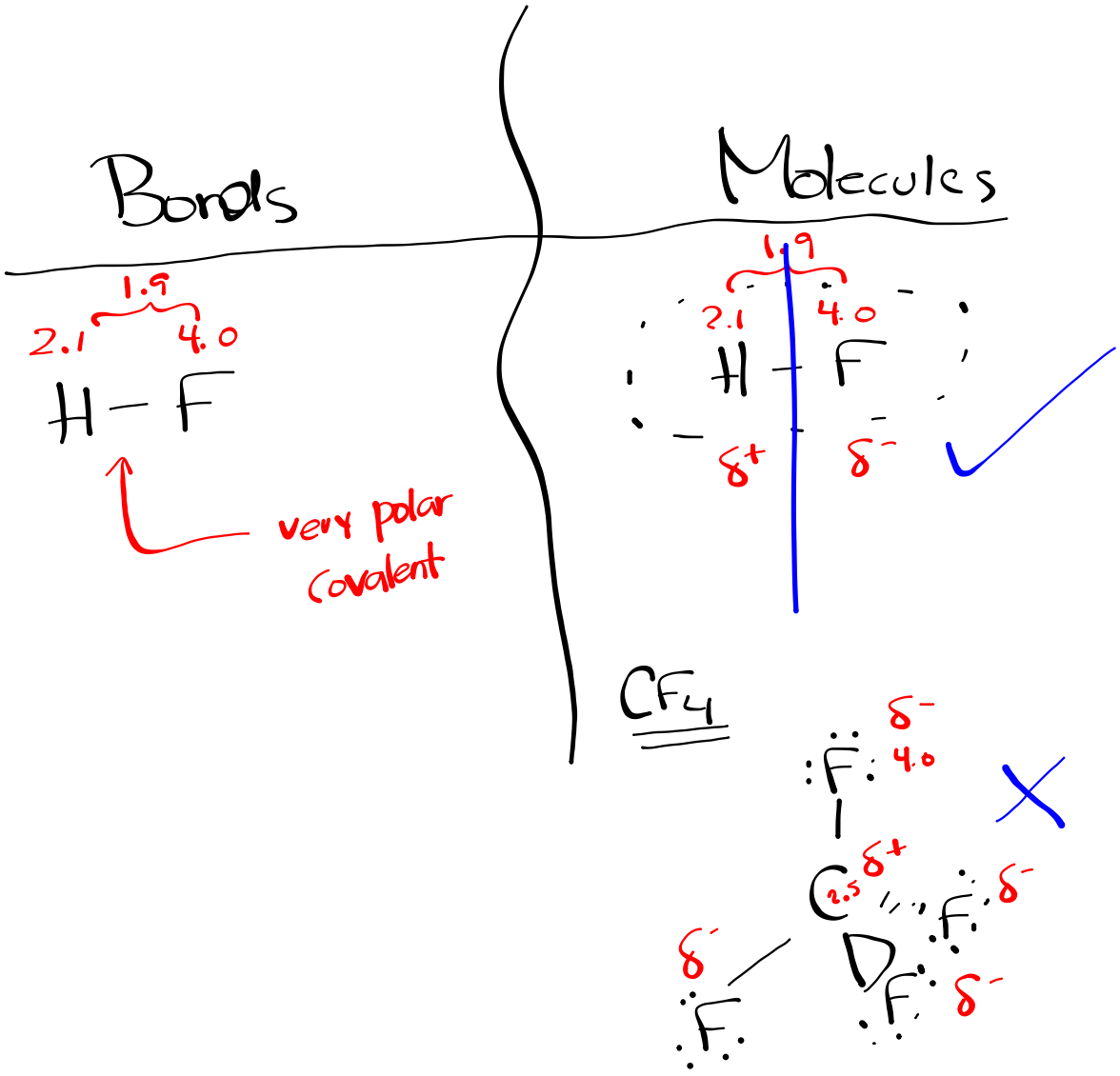
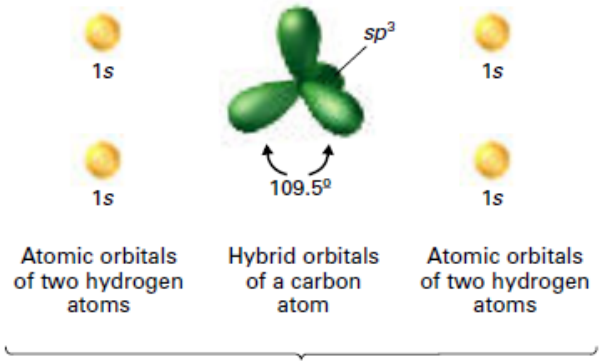
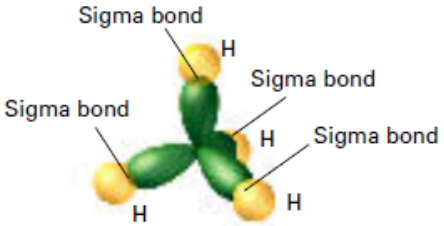

Bonds
Molecules
1.9
2.1
4.0
H – F
very polar covalent
δ+
δ−
CF4
2.5
X

Table 6.2

Electronegativity Values for Selected Elements

H 2.1						
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9		





Methane molecule

Attraction Between Molecules

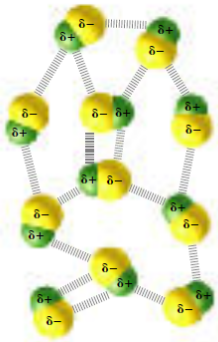
Intermolecular forces are weaker than both ionic and covalent bonds.

Van der Waals Forces

-Weakest attractions between molecules.
-Can be separated into two categories:

Dipole Interactions

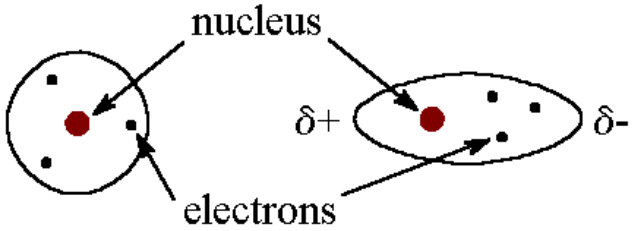
Electrical attraction between oppositely charged regions of polar molecules.



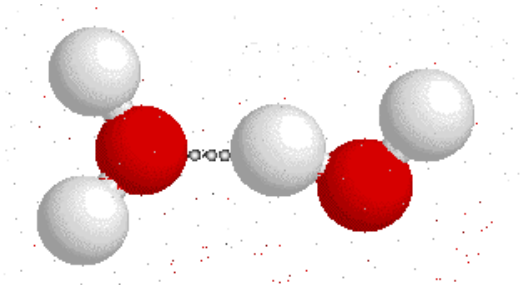
Dispersion Forces (London Dispersion Forces)

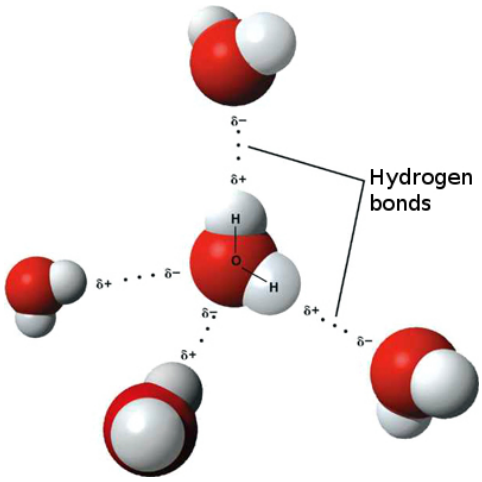
-weakest of all molecular interactions
-occur between even non-polar molecules
-caused by the motion of electrons

when moving electrons are momentarily on one side of a molecule, the electrons of the neighbouring molecule will move t the opposite side, causing a weak attraction.



Hydrogen Bonds





Hydrogen Bonds

Strong attractive forces in which a hydrogen covalently bonded to a very electronegative atom (O, N, F), is weakly bonded to an unshared electron pair of another electronegative atom.

- strongest intermolecular force
- not as strong as an ionic or covalent bond