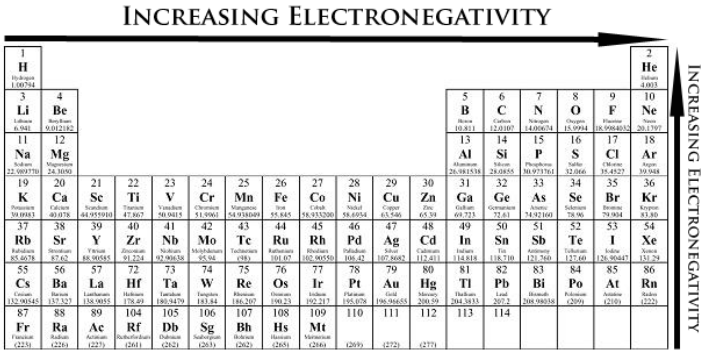
CHAPTER 1
INTRODUCTION TO ORGANIC CHEMISTRY

1.1 Review of chemical bonding fundamentals

Atoms try to achieve an s^2p^6 electron configuration (full outer valence shell, stable octet).

They can do this by forming
a) covalent bonds (sharing electrons equally),
b) polar covalent bonds (sharing electrons unequally) or
c) ionic bonds (losing or gaining electrons).

One can determine the type of bond formed between 2 atoms by comparing the electronegativity values of the 2 atoms. **Electronegativity** is a number between 1.0 and 4.0 that indicates how strongly an atom attracts electrons. Atoms on the left hand side of the periodic table have low values of electronegativity (less than 2.0) and atoms on the right hand side of the periodic table have higher values of electronegativity.



A difference more than or equal to 2.0 is an ionic bond
A difference between 0.4 and 2.0 is a polar covalent bond.
A difference in electronegativity less than or equal to 0.4 is a covalent bond.

The electronegativity of the atoms we will be dealing with this term:

H = 2.2; C = 2.6; O = 3.5; N = 3.1; F = 4.0; Cl = 3.2; S = 2.6; Br = 2.9; I=2.7

KNOW:
Covalent (non-polar): C-C, C-H, C-S, S-H
Moderately polar: C-Cl
Polar: O-H, N-H, C-N, C-O

Bonding in molecules is conveniently shown using Lewis dot diagrams which show the number of electrons in their outer shell.

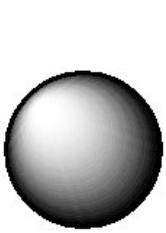
IA	IIA	IIIA	IVA	VA	VIA	VIIA	Noble Gases
H ·							**He :**
Li ·	**· Be ·**	**· B ·**	**· C ·**	**: N ·**	**: O ·**	**: F :**	**: Ne :**
Na ·	**· Mg ·**	**· Al ·**	**· Si ·**	**: P ·**	**: S ·**	**: Cl :**	**: Ar :**
K ·	**· Ca ·**					: Br :	**: Kr :**
Rb ·	**· Sr ·**					**: I :**	**: Xe :**
Cs ·	**· Ba ·**						

Know the Lewis dot structures of the highlighted atoms.

Bonding examples using Lewis dot:

2 F atoms : F · + · F : ⟶ : F : F : or : F - F :

Review of s and p orbital shapes:



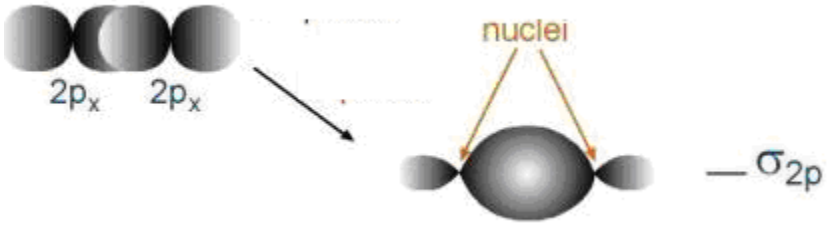
s orbital p orbital

Orbital overlap view:

F (atomic # 9) $1s^2 2s^2 2p^5$

F $1s^2 2s^2 2p^5$

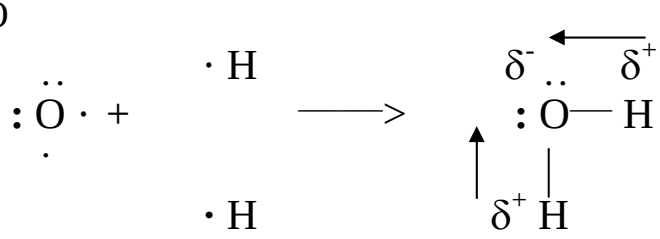
Each F atom is one electron short of a full outer shell. It can fill that unfilled p orbital by overlapping the unpaired electron of one F atom with the unpaired electron (with opposite spin) of a second F atom.



The resulting bond has its electron density concentrated on the line between the two F nuclei and is called a sigma bond.

Lewis dot notation for molecules containing O atoms:

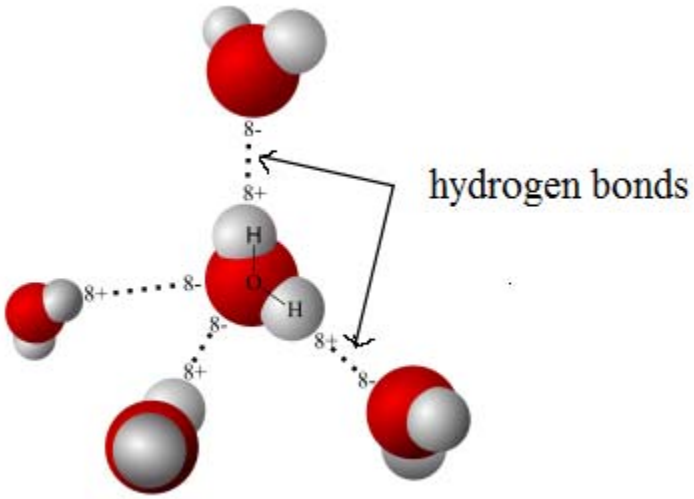
a)H_2O



The arrows indicate the dipoles: the direction in which the highest shared electron density occurs in the bond.

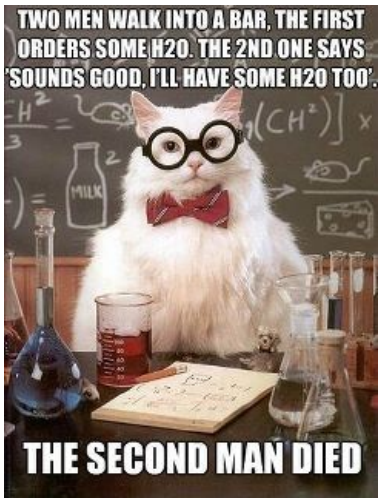
There are particularly strong attractions between the δ^- on the O of one water molecule and the δ^+ on the H atom of another water molecule. These attractions are called **hydrogen bonds.** They are very important attractions not only in water but in proteins,

carbohydrates, and nucleic acids. Hydrogen bonds are a specific type of polar attraction between a H with a δ^+ and a very electronegative atom F, O, or N.



Modified from Wikipedia

b) Do the Lewis dot diagram for H_2O_2.



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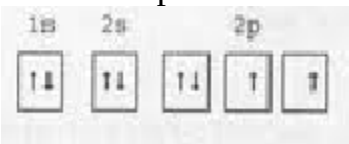
b) O_2

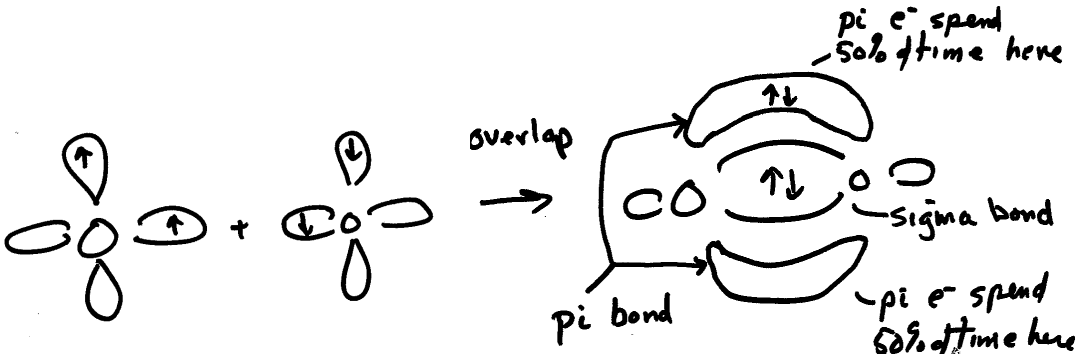
$$:\ddot{O}\cdot \quad + \quad \cdot\ddot{O}: \quad \longrightarrow \quad \ddot{O}{=}\ddot{O}$$

. . .. ..

The result is a double covalent bond with 4 shared electrons (2 pairs).

Orbital overlap representation of O_2 :
O: $1s^2 2s^2 2p^4$



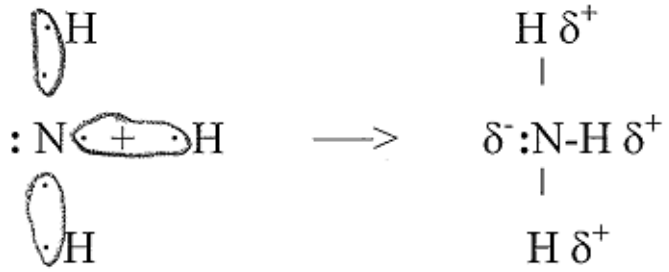


The 2 electrons in the pi orbital are not as close to the two positively charged O nuclei and can more easily react with another atom or especially a positively charged ion. In general pi electrons of a double bond will be more reactive than the sigma bond.

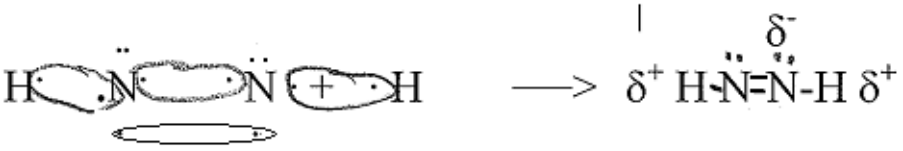
In both molecules above O atom forms two bonds, either two single bonds or one double bond

Lewis dot structures of molecules with N atoms

a)ammonia, NH_3

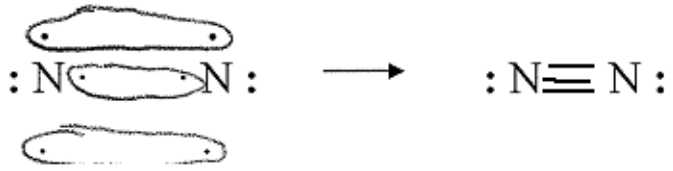


b) N_2H_2



In this molecule each N atom forms a double and a single bond.

c) N_2



This produces a triple covalent bond with 6 shared electrons (3 pairs).

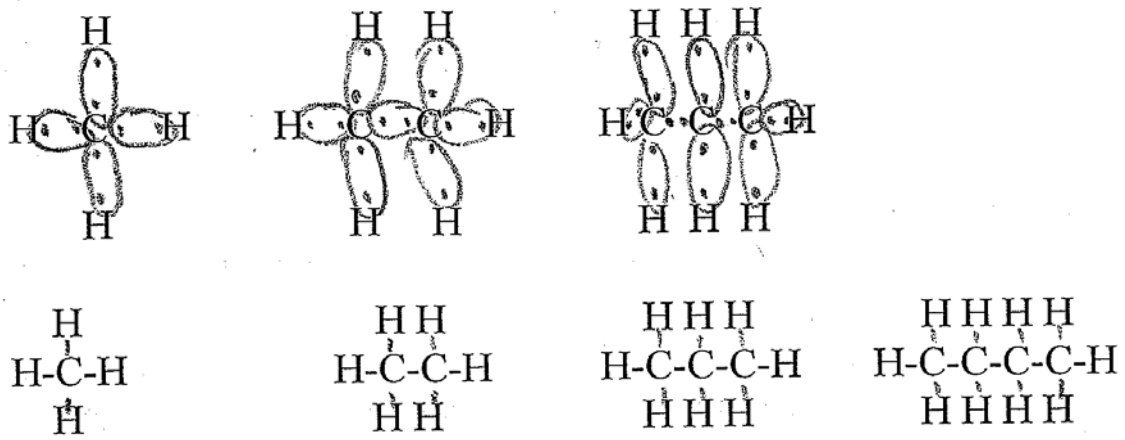
In all of the above molecules the 3 unpaired electrons of N form 3 bonds: 3 single bonds, a double and a single, or a triple bond.

What will the orbital overlap picture of N_2 look like?

Lewis dot structures for molecules with C atoms

C is in Group IV and it forms 4 bonds:

CH_4 C_2H_6 C_3H_8 C_4H_{10}



Since H is the most common atom found bonding with C we will sometimes simply show 4 bonds coming off of a C without explicitly showing the atom at the end of the bond. When we do this, it is assumed that H is the atom bonded. If some other atom beside H is bonded to C, we HAVE to show it explicitly.

These chains can be almost infinitely long and these long chains are what make the chemistry of carbon so extensive and so diverse.