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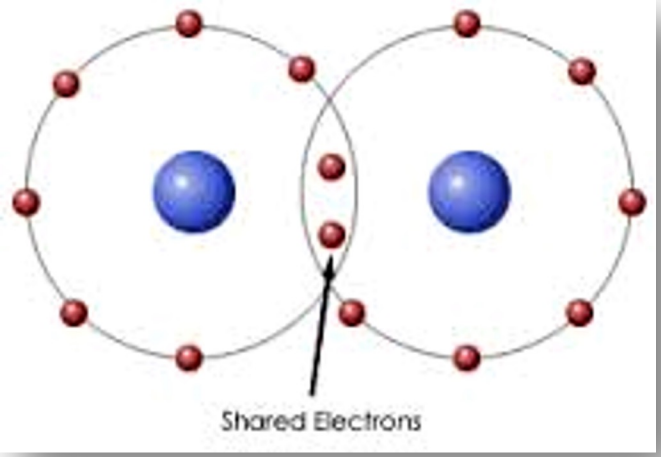
Chapter 3

Chemical Bonds II

Lecture 6



Chemical Bonds II

OUTLINE

- Molecular compounds
- Electronegativity
- Polarity of bonds
- Lewis structure

Objectives

- ***After you have studied this chapter, you should be able to***
- Use the basic vocabulary of Electronegativity
- Distinguish between polar and nonpolar compounds
- Learn how to write Lewis structure of some compounds

Molecular Compounds

- **Molecular compound:** Only covalent bonds
- Naming molecular compounds
  - the less electronegative element is named first (it is generally written first in the formula)
  - prefixes "di-", tri-", etc. are used to show the number of atoms of each element; the prefix "mono-" is generally omitted
    - NO is nitrogen oxide (nitric oxide)
    - SF_2 is sulfur difluoride
    - N_2O is dinitrogen oxide (laughing gas)
    - Exception: carbon monoxide

Mono

di

Tri

Tetra

Penta

hexa

CO

CO_2

SO_2

SO_3

H_2O

P_2O_5

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Electronegativity

Table 3.5 Electronegativity Values for the Elements (Pauling Scale)

1A	2A											3A	4A	5A	6A	7A
							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	3B	4B	5B	6B	7B	8B			1B	2B	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2

<1.0 | 1.0–1.4 | 1.5–1.9 | 2.0–2.4 | 2.5–2.9 | 3.0–4.0

F has highest value Noble gases have 0 value

Ionic bonds form when electronegativity difference ≥ 1.9

NaCl — Ionic

$3 - 0.9 = 2.1 > 1.9$

HCl $= 3 - 2.1 = 0.9 > 0.4$

Polar

O_2 $3.5 - 3.5 = 0$

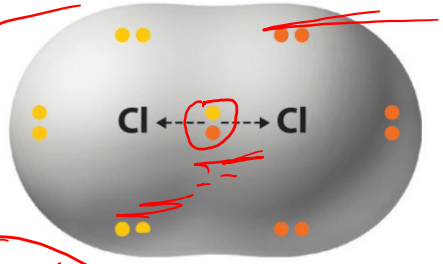
non-polar

Polarity of Bonds

Nonpolar: Electrons are shared equally

Polar: Electrons are NOT shared equally

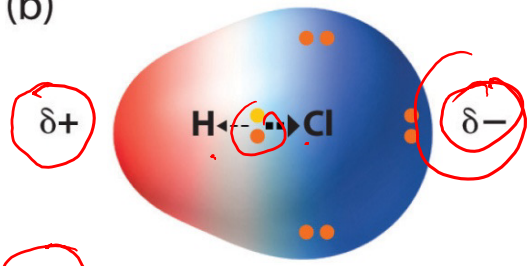
(a)



Nonpolar covalent bond

Bonding electrons shared equally between two atoms.
No charges on atoms.

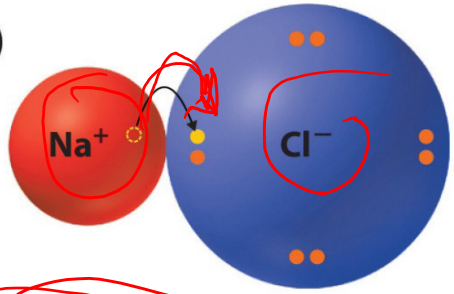
(b)



Polar covalent bond

Bonding electrons shared unequally between two atoms.
Partial charges on atoms.

(c)



Ionic bond

Complete transfer of one or more valence electrons.
Full charges on resulting ions.

Electronegativity difference (Δ)	Bond
$\Delta > 2$	Ionic
$0.4 < \Delta < 2$	Polar Covalent, (or covalent with partial ionic character).
$\Delta < 0.4$	Covalent

non-polar

Polarity of Covalent Bonds

Bond	Difference in Electronegativity	Type of Bond
H-Cl	3.0 - 2.1 = 0.9	polar covalent
O-H	3.5 - 2.1 = 1.4	polar covalent
N-H	3.0 - 2.1 = 0.9	polar covalent
Na-F	4.0 - 0.9 = 3.1	ionic
C-Mg	2.5 - 1.2 = 1.3	polar covalent
C-S	2.5 - 2.5 = 0.0	nonpolar covalent

CO_2

O_2

F O Cl N

Br

HF

CH_4

non-polar

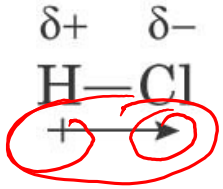
CH_3Cl

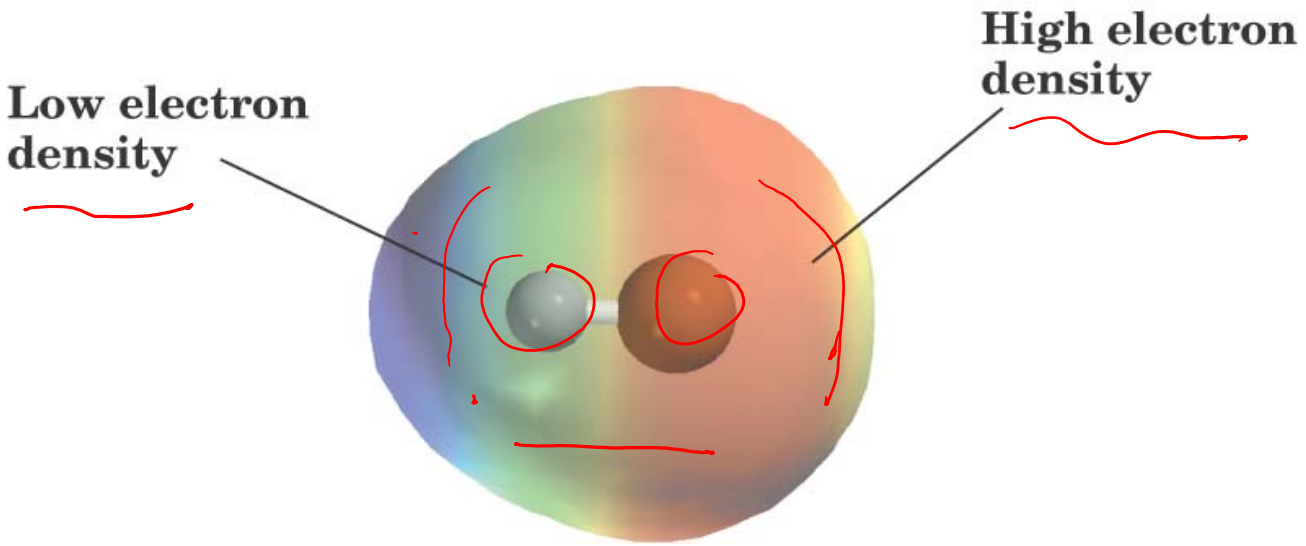
Polar

metal + non-metal

Ionic

Polarity of Covalent Bonds



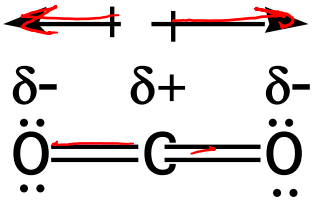


More electron density shown by δ^- or the head of a crossed arrow

Less electron density shown by δ^+ or the tail of a crossed arrow

Polarity of Molecules

- Polar molecule has
  - Polar bonds, and
  - Partial positive and partial negative charges in different parts of molecule, i.e., is a dipole (has two poles)
- Carbon dioxide, CO_2, has two polar bonds but, because of its geometry, the pulls balance out so it is a nonpolar molecule

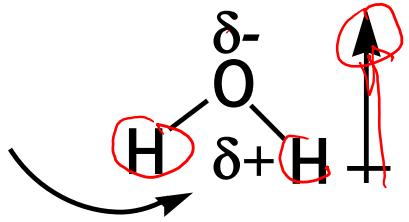


Carbon dioxide
(a nonpolar molecule)

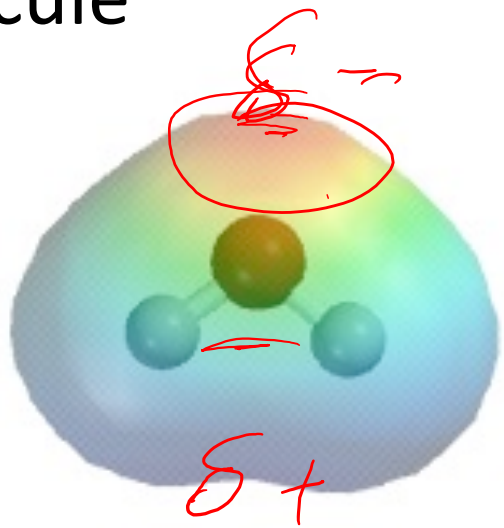
Polarity of Molecules

Water, H_2O, has two polar bonds and, because of its geometry, is a polar molecule

center of partial positive charge is midway between the two hydrogen atoms

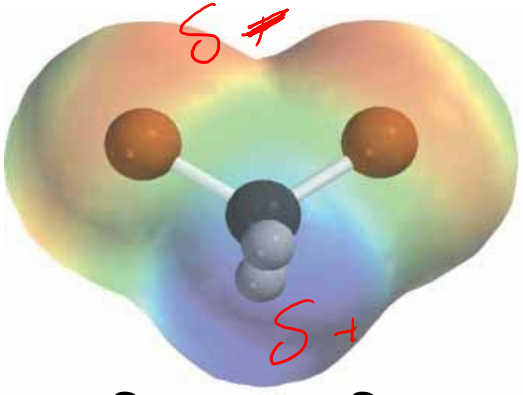


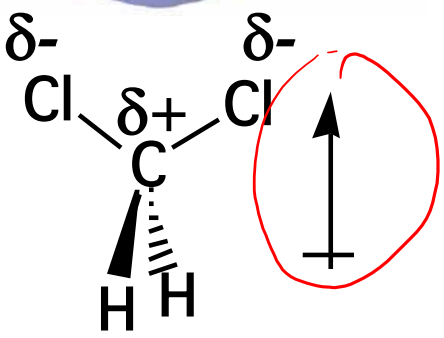
Water
(a polar molecule)



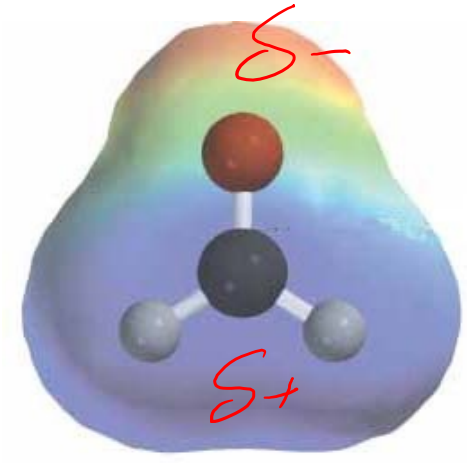
Polarity of Molecules

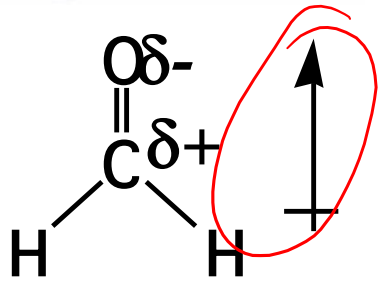
- Both dichloromethane, CH_2Cl_2, and formaldehyde, CH_2O, have polar bonds and are polar molecules



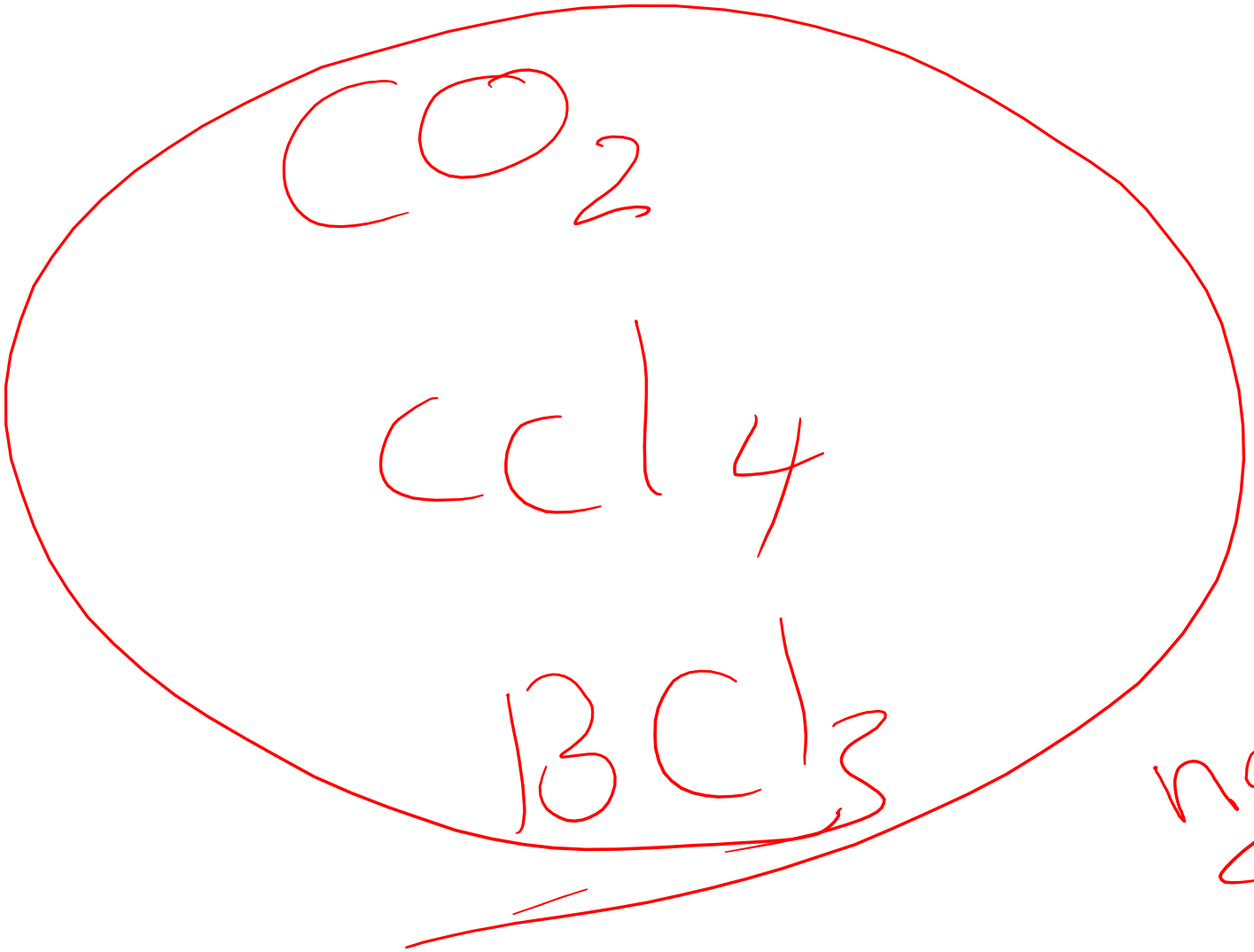


Dichloromethane





Formaldehyde



non-Polar

Lewis Structures

- ✓ Used to decide on the arrangement of atoms in the molecule
- ✓ Bonding (shared) electrons are shown as bonds (lines)
- ✓ Nonbonding electrons are represented as a pair of Lewis dots

periodic table 2.swf

Drawing Lewis Structures

1. Determine the number of valence electrons in the molecule
2. Decide on the arrangement of atoms in the molecule
3. Connect the atoms by single bonds
4. Show bonding electrons as a single line; show nonbonding electrons as a pair of Lewis dots
5. In a single bond, atoms share one pair of electrons; in a double bond, they share two pairs, and in a triple bond they share three pairs.

H_2O $2 \times 1 + 6 = 8\ e^- = 4$ pair

(2)

CO_2 $4 + 2 \times 6 = \frac{16\ e^-}{2} = 8$ pair

CCl_4

Lewis Structures

- Examples: (the number of valence electrons is given in parentheses after the molecular formula

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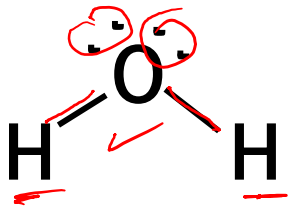
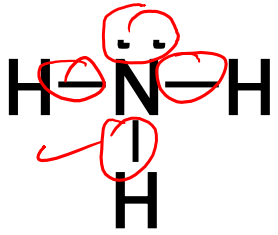


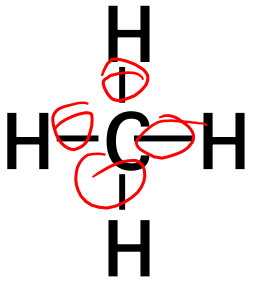
H_2O (8)

Water



NH_3 (8)

Ammonia



CH_4 (8)

Methane

$3 \times 1 + 5 = 8$

Exceptions to the Octet Rule

- ✓ H and He have a maximum of 2 electrons (duet)
- ✓ Period 2 elements have a maximum of 8 electrons (use 2s and 2p orbitals)
- ✓ Atoms of period 3 elements may have more than 8 electrons

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End of the Lecture

مع تحيات أ/ عبيده محمد

دعواتي للجميع بالتوفيق والنجاح