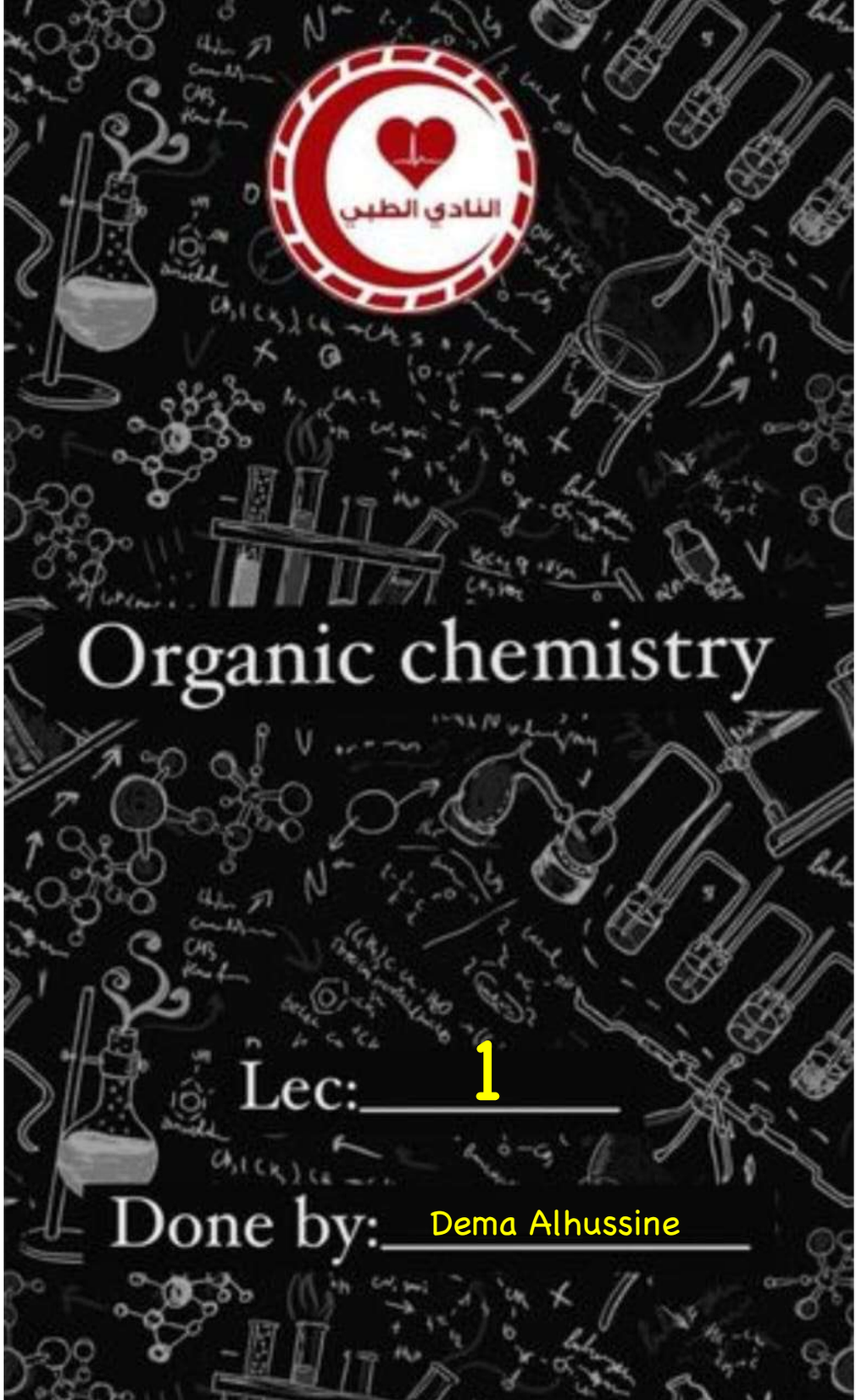


Organic chemistry

Lec: 1

Done by: Dema Alhussine

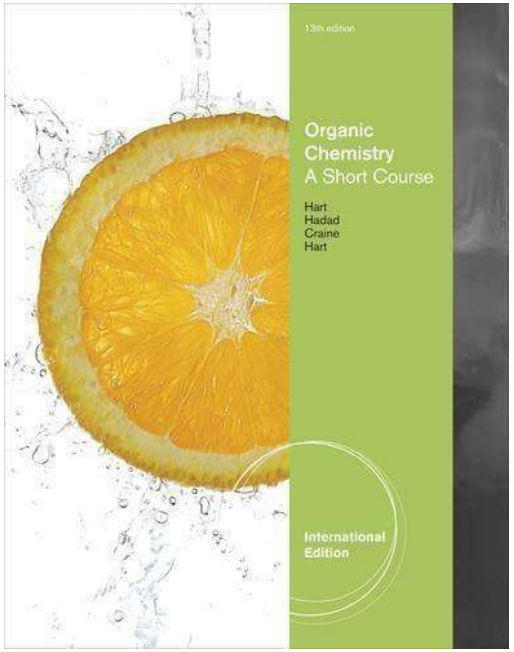
Chem 237 Basics of Organic Medicinal Chemistry

- **Course description**

This is the first year organic chemistry course, introducing basic concepts and principles of organic chemistry (chapters 1 – 11).

- **Texts**

Hart, Craine, Hart and Hadad, Organic Chemistry, A Short Course,13th Edition (Brooks/Cole, Cengage Learning, CA 94002-3098 USA, 2012).



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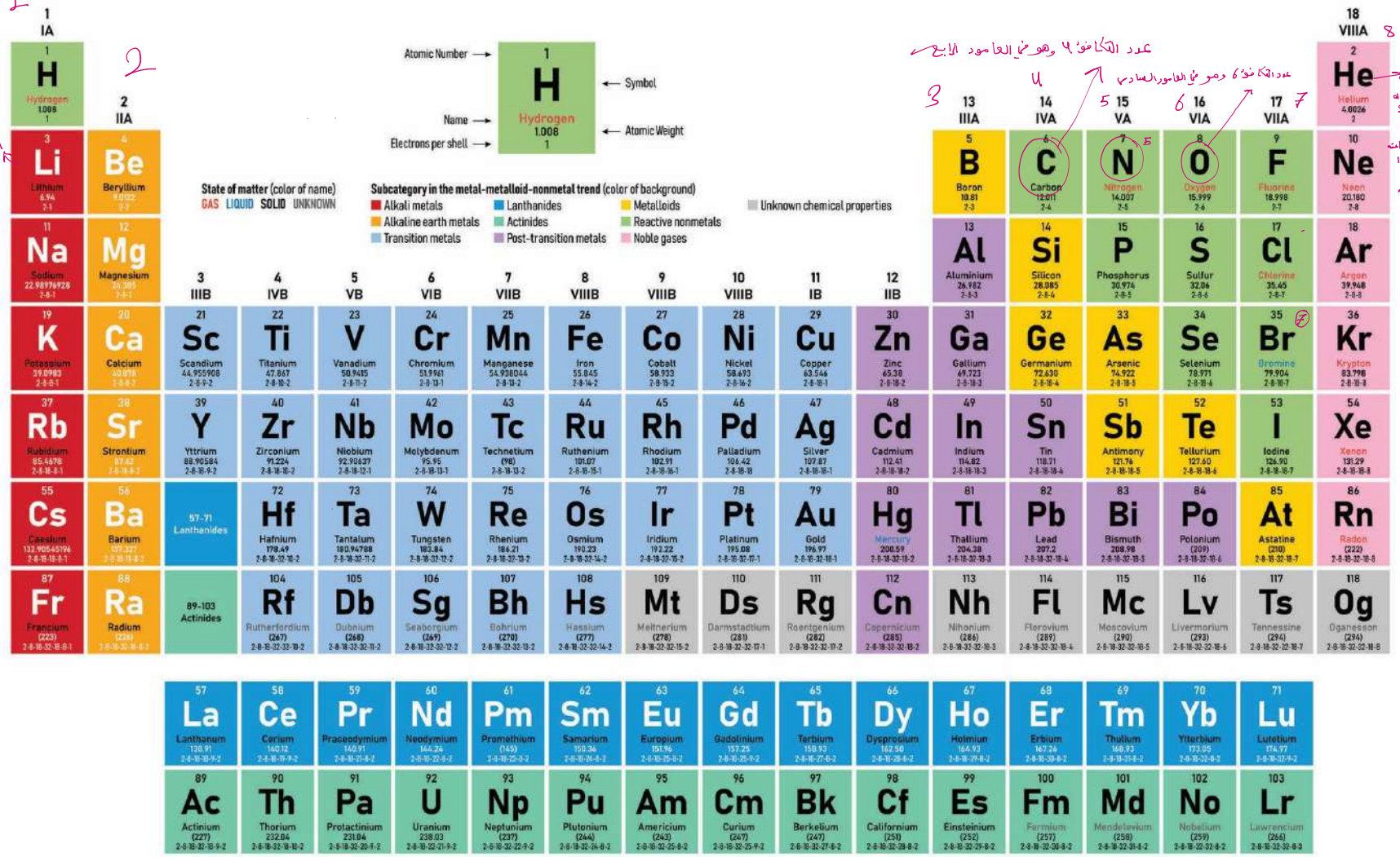
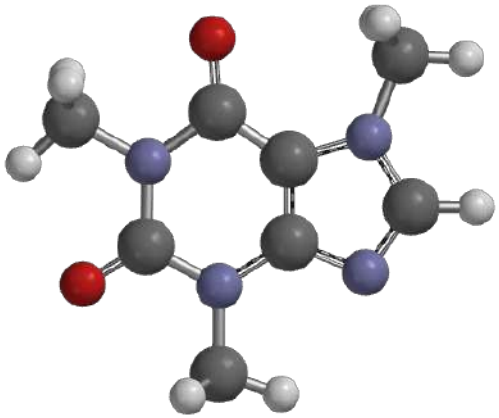
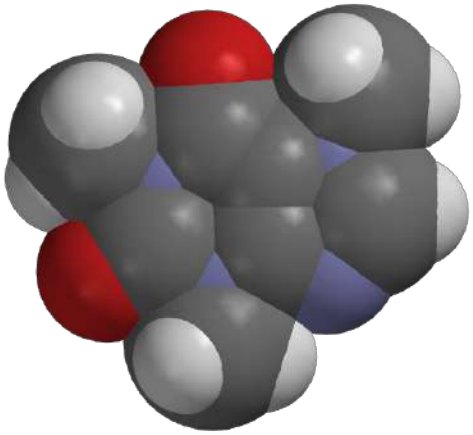
VE = group number
Valence electrons
(Column)

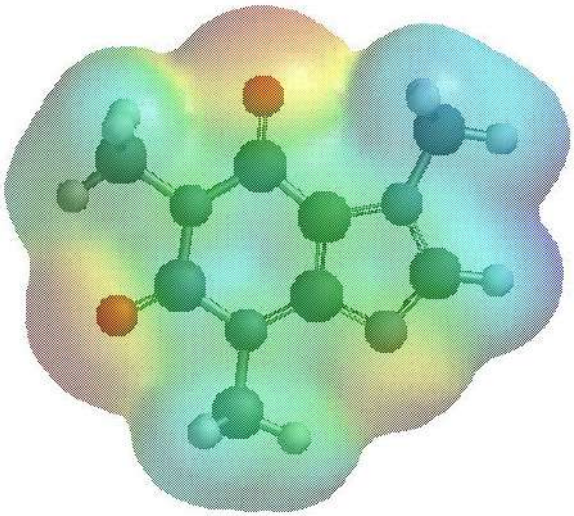
Periodic Table of the Elements







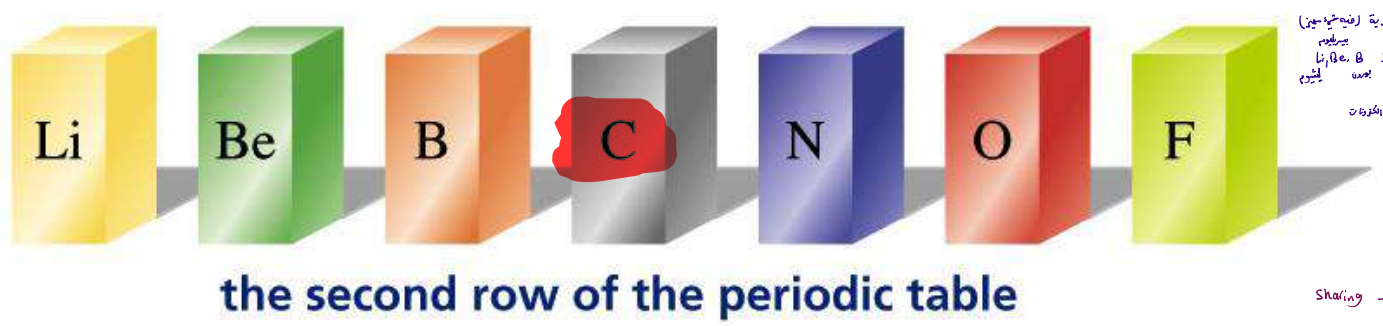
Chapter 1: Bonding and Isomerism



Organic Chemistry

- Organic compounds are compounds containing carbon

major element



- Atoms to the left of carbon give up electrons.
- Atoms to the right of carbon accept electrons.
- Carbon shares electrons.

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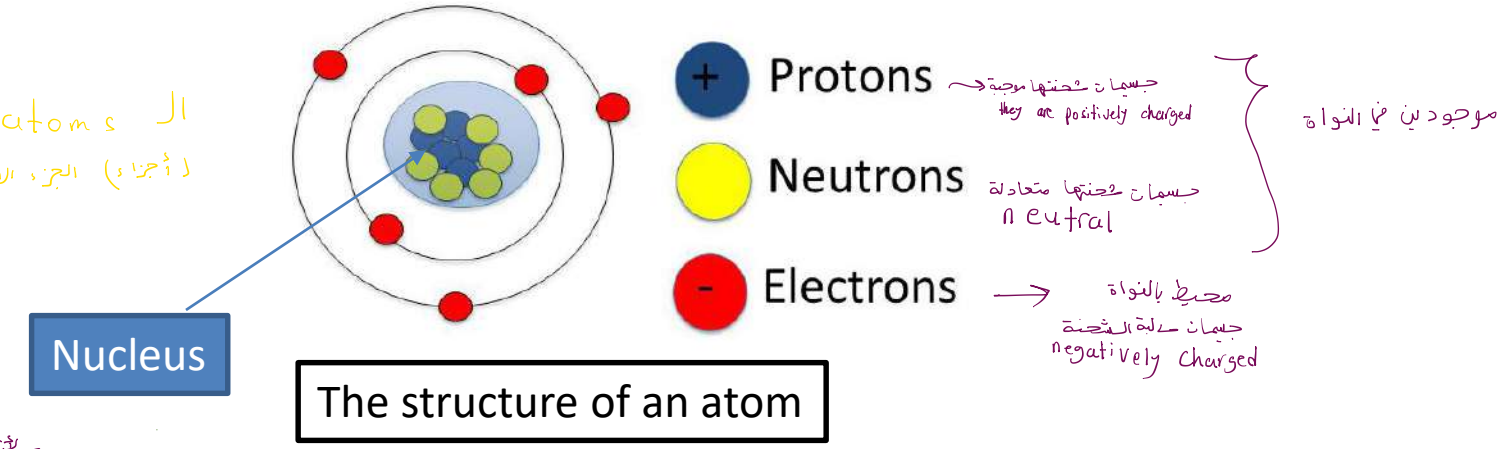
Bonding and Isomerism

1.1 How Electrons Are Arranged in Atoms

- An atom is: the *smallest particle* of an element that retains all of the chemical properties of that element.

- An atom consists of negatively charged electrons, positively charged protons, and neutral neutrons



The structure of an atom

- **Atomic number**: numbers of protons in its nucleus and it's the number of electrons in the neutral atom.

- **Mass number**: the sum of the protons and neutrons of an atom.

(Protons and neutrons are ~1837 times the mass of an e^-)

- Isotopes have the same atomic number but different mass numbers (^{12}C and ^{13}C)

- Electrons are located in atomic orbitals (S, P, d, f).
- Orbitals tell us the energy of the electron and the volume of space around the nucleus where an electron is most likely to be found.
- Orbitals are grouped in shells .

Each orbital can hold a maximum of 2e⁻ and the two electrons have opposite spin

Table 1.1 Distribution of Electrons in the First Four Shells That Surround the Nucleus

	First shell	Second shell	Third shell	Fourth shell
Atomic orbitals	*s*	*s*, *p*	*s*, *p*, *d*	*s*, *p*, *d*, *f*
Number of atomic orbitals	1	1, 3	1, 3, 5	1, 3, 5, 7
Maximum number of electrons	2	8	18	32

Example :

^{1}H: $1s^1$

^{6}C : $1s^2\ 2s^2\ 2p^2$

Valence electrons (VE) are located in the outermost shell. They are involved in chemical reactions.

VE = Group number

Examples:	^{1}H: Is^1	1
	^{8}O: $IS^2\ 2S^2\ 2P^4$	6
	^{6}C: _____	

VE

Lewis symbol of atom

H·

O:

·C·

Table 1.3 Valence Electrons of the First 18 Elements

Group	I	II	III	IV	V	VI	VII	VIII
	H·							He:
	Li·	Be·	·B·	·C·	·N:	·O:	:F:	:Ne:
	Na·	Mg·	·Al·	·Si·	·P:	·S:	:Cl:	:Ar:

- **Chemical Bonds**

1. Ionic Bonding

A – B

Complete transfer of electrons from one element to another

An ionic bond is an electrostatic attraction between positive & negative ions resulting from e⁻ transfer.

$^{11}Na : [Ne]\,3s^1 \xrightarrow{-e} Na^{\oplus} : [Ne]$

$^{17}Cl: [Ne]\,3s^2\,3p^5 \xrightarrow{+e} Cl^{\ominus} : [Ar]$

$1s^2\ 2s^2\ 2p^6$

$Na^{+} + Cl^{-} \rightarrow Na\text{-}Cl$ (ionic bond)

The resulting e⁻ configuration of both ions are those of the nearest noble gas, Ne and Ar respectively, both satisfy the octet rule.

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2. Covalent Bonding

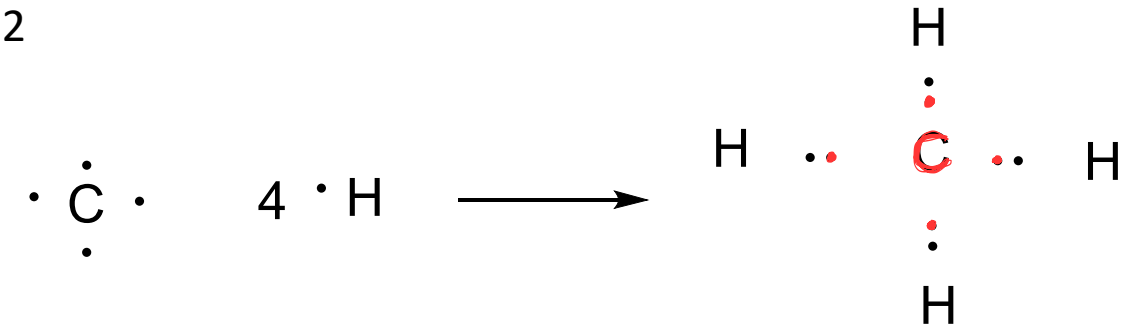
- Ionic bonds occur when an e^- is transferred between a metal and nonmetal.
- Covalent bonds are resulting from sharing e^-

$$2\,H\cdot \longrightarrow H{:}H$$

The result is both atoms have a [He] e^- configuration, *i.e.*

The bond is commonly display as a line rather than a pair of e^- (:), *i.e.* H - H rather than H : H

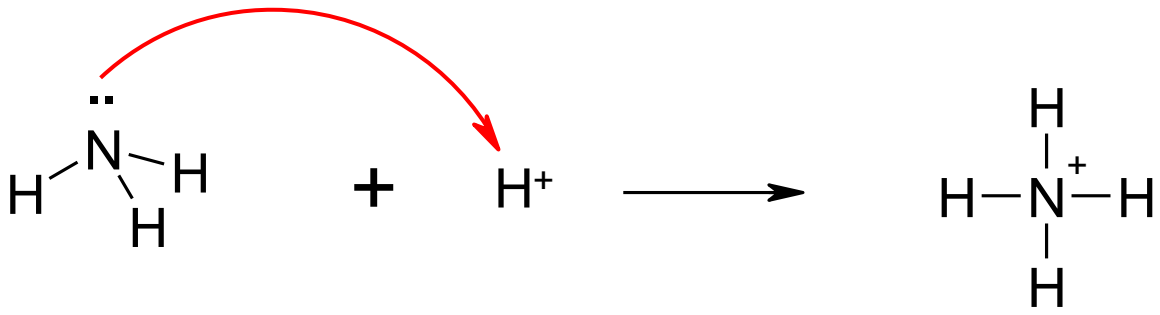
Example 2



A second general version of a covalent bond is possible. This occurs when BOTH e^- come from one atom: a coordinate covalent bond

i.e.

$$NH_3 + H^+ \rightarrow NH_4^+$$



Electronegativity (EN) : measures the tendency of an atom to attract a shared pair of electrons (or electron density).

A — B (covalent bond)

عنصر يستطيع أن يجعل e قريب منه أكثر من العنصر الآخر

TABLE 1.3 The Electronegativities of Selected Elements[a]

IA	IIA	IB	IIB	IIIA	IVA	VA	VIA	VIIA
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							Br 2.8
								I 2.5

increasing electronegativity →

increasing electronegativity ↑

[a]Electronegativity values are relative, not absolute. As a result, there are several scales of electronegativities. The electronegativities listed here are from the scale devised by Linus Pauling.

* يوجد نوعين من Covalent bonds
  1. Non polar
  2. Polar Covalent bond

C → 2.5

H → 2.1

* من اليسار إلى اليمين EN بتزيد
* طالعة طلوع EN بتزيد

مثلا لو بدي أعمل مقارنة بين ال

Covalent bonds can be classified as

A. Nonpolar covalent bond (Δ EN = 0-0.5) → إذا كان الفرق في EN صغير

Examples C-C C-H

B. Polar covalent bond (Δ EN = 0.5-1.9)

A polar bond has a negative end and a positive end

dipole moment (D) = μ = *e* x *d*

(*e*) : magnitude of the charge on the atom

(*d*) : distance between the two charges

Table 1.4 The Dipole Moments of Some Commonly Encountered Bonds

Bond	Dipole moment (D)	Bond	Dipole moment (D)
H—C	0.4	C—C	0
H—N	1.3	C—N	0.2
H—O	1.5	C—O	0.7
H—F	1.7	C—F	1.6
H—Cl	1.1	C—Cl	1.5
H—Br	0.8	C—Br	1.4
H—I	0.4	C—I	1.2

if The Δ EN increases the polarity increases

Note : If Δ EN is more than 1.9then the bond is ionic Ex: Li-F

Bond Polarity & Electronegativity (cont'd)

The result of polar covalent bonding is that the e^- pair spend more time near the more EN atom. This means it will acquire a permanent excess negative charge. The other atom acquires a permanent excess positive charge. This is indicated by a δ^+ or δ^- (where δ means a "partial charge") or a dipole arrow which points from the positive end of the bond to the negative end.

$\delta^+ \quad \delta^-$

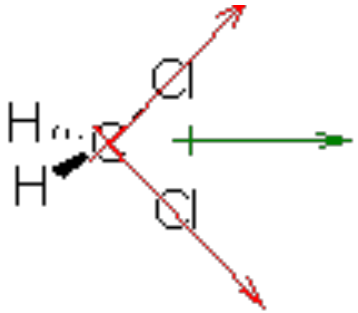
H—Cl

polar bond

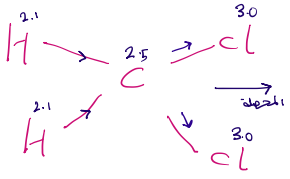
H—Cl

Bond Polarity & Electronegativity (cont'd)

The more polar the molecule the stronger the dipole moment. The molecular dipole moment is the vector sum of the bond moments, *i.e.*

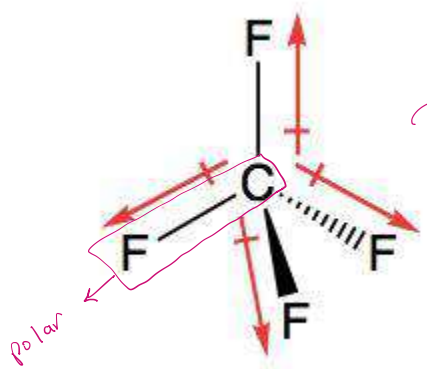


Net dipole moment

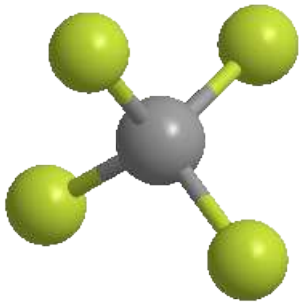


$N \neq 0$

* الأرقام ليست للحفظ



non polar molecule



Net dipole moment =0

ولكن بهما تحريجات تهضا المحصلة لتحديد إذا ما كان Polar أو لا

28-2-2024 Wednesday