

(8) Today

Sections 2.1 - 2.4
Polar Covalent Bonds, Formal Charges,
Resonance/Electron Delocalization

Section 2.1: Problems 2-1 – 2-4
Section 2.2: Problems 2-5, 2-6, 2-28 – 2-30,
2-32, 2-53, 2-59, 2-63
Section 2.3: Problems 2-7, 2-8, 2-35, 2-36

(10) Second Class from Today

Sections 2.7 – 2.11
Acids and Bases

Next Class (9)

Sections 2.3 – 2.7
Formal Charges, Resonance/Electron
Delocalization

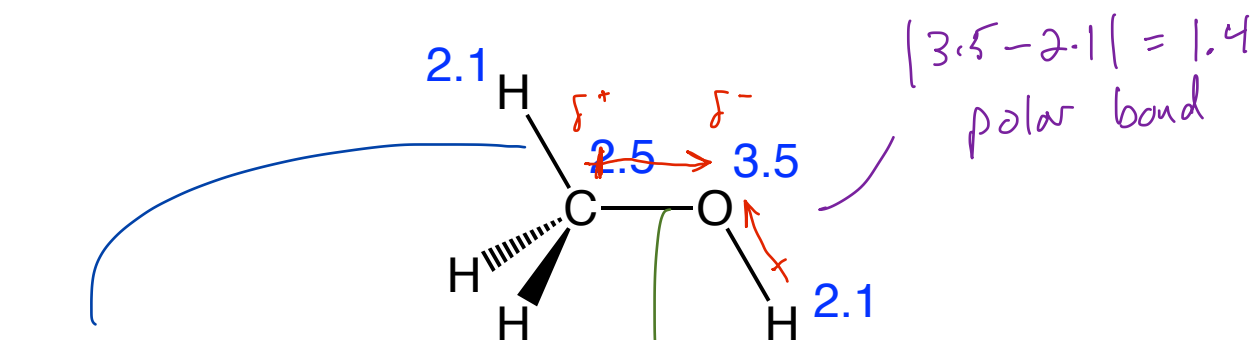
Sections 2.4 – 2.6: Problem 2-9, 2-10, 2-20,
2-21, 2-23, 2-26, 2-37, 2-38, 2-56, 2-57,
2-61,
Challenging Problems 2-33, 2-34, 2-39

Third Class from Today (11)

Sections 2.7 – 2.11
Acids and Bases

Section 2-12
Non-Covalent Interactions Between
Molecules

"As a rough guide, bonds between atoms whose electronegativities differ by less than 0.5 are [considered] nonpolar covalent, bonds between atoms whose electronegativities differ by 0.5 to 2 are polar covalent, and bonds between atoms whose electronegativities differ by more than 2 are largely ionic."¹

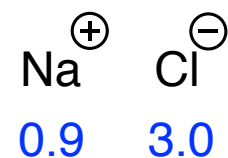


$$|2.1 - 2.5| = 0.4$$

C-H bond is considered nonpolar even though the H is slightly δ^+ and the C is slightly δ^-

$$|3.5 - 2.5| = 1.0$$

polar bond



$$|3.0 - 0.9| = 2.1$$

ionic



¹ Organic Chemistry, 10th ed., McMurray, OpenStax (2023) <https://openstax.org/details/books/organic-chemistry>

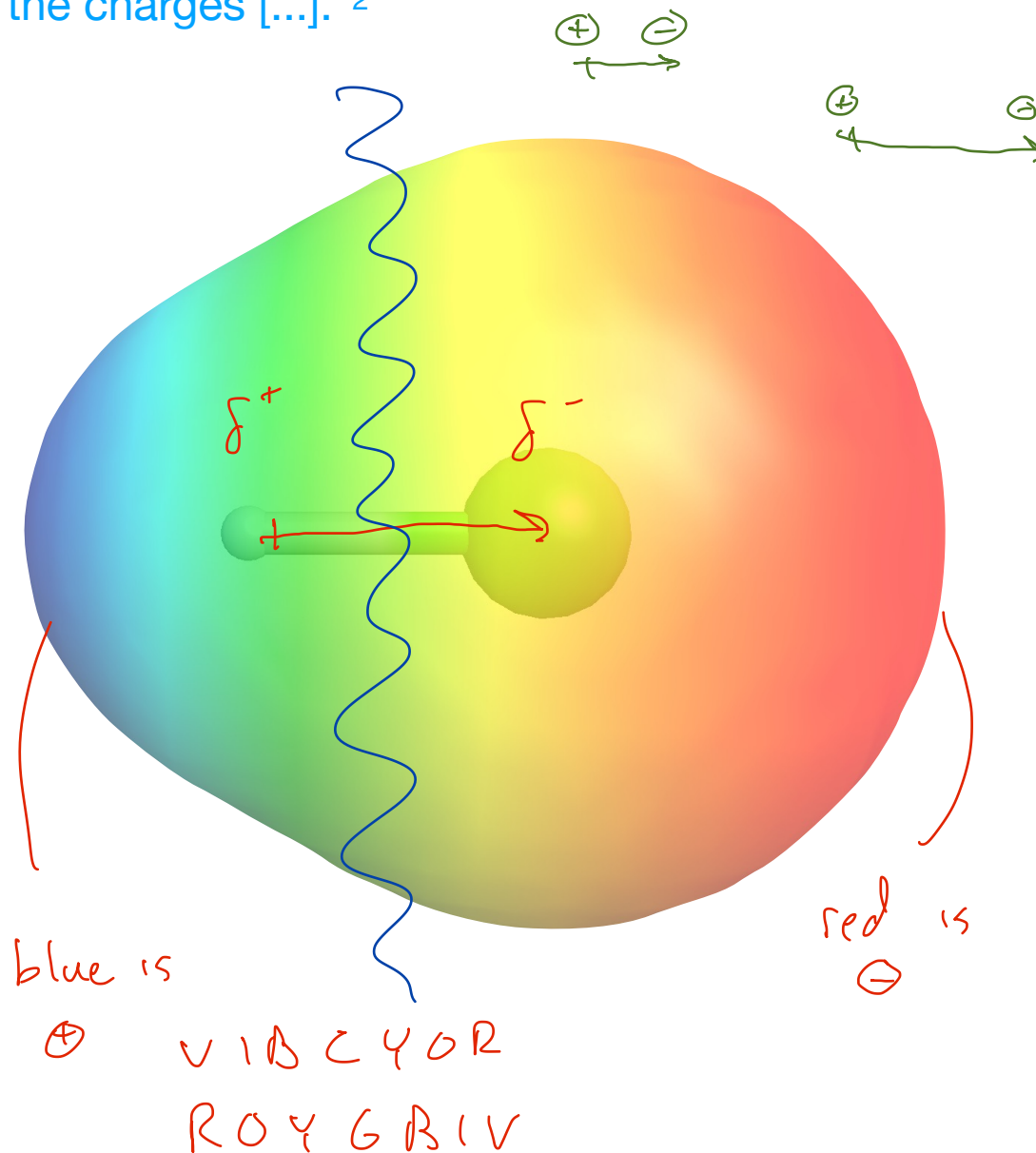
Polar Covalent Bonds and Dipole Moments: HCl \leftarrow 3.0 $|3.0 - 2.1| = .9$
 \uparrow 2.1

"The dipole moment, μ [...], is defined as the magnitude of the charge Q at either end of the [...] dipole times the distance r between the charges [...]."²

H-Cl is a polar molecule
 the \oplus and \ominus are
 on opposite sides
 of the molecule

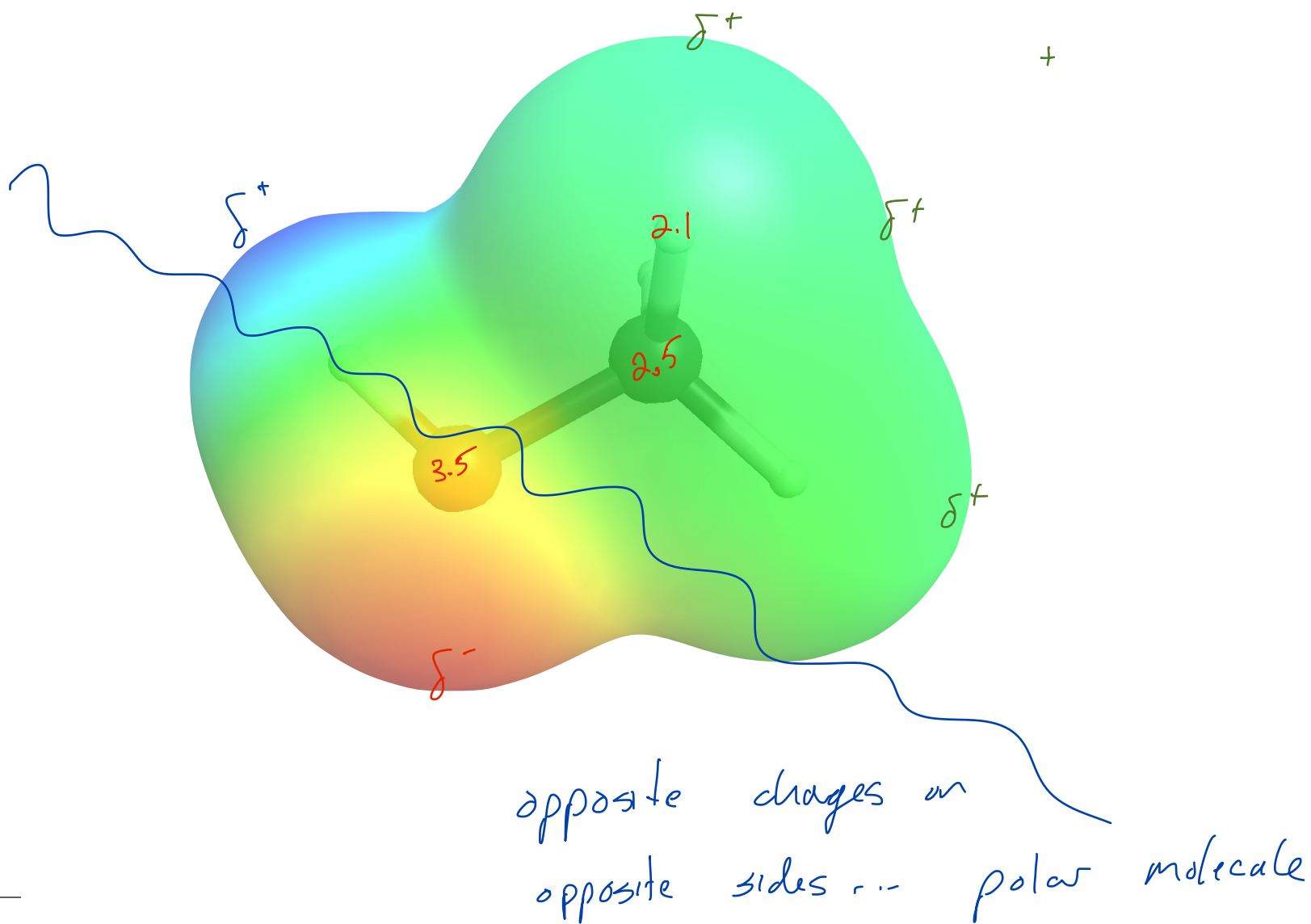
Larger energy differences...
 stronger dipoles

Larger separation of charge
 stronger dipoles



² Ibid.

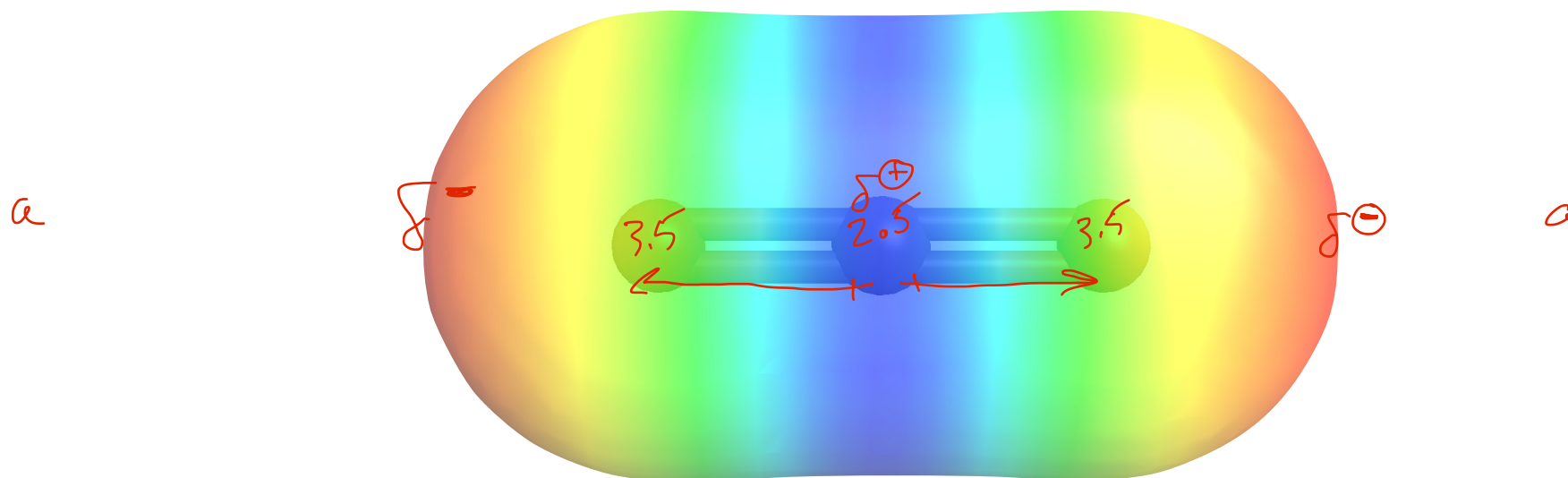
"The dipole moment, μ [...], is defined as the magnitude of the charge Q at either end of the [...] dipole times the distance r between the charges [...]." ³



³ Ibid.

"The dipole moment, μ [...], is defined as the magnitude of the charge Q at either end of the [...] dipole times the distance r between the charges [...]." ⁴

b



opposite charges cannot be separated by drawing
a line (plane) through the molecule
b
nonpolar

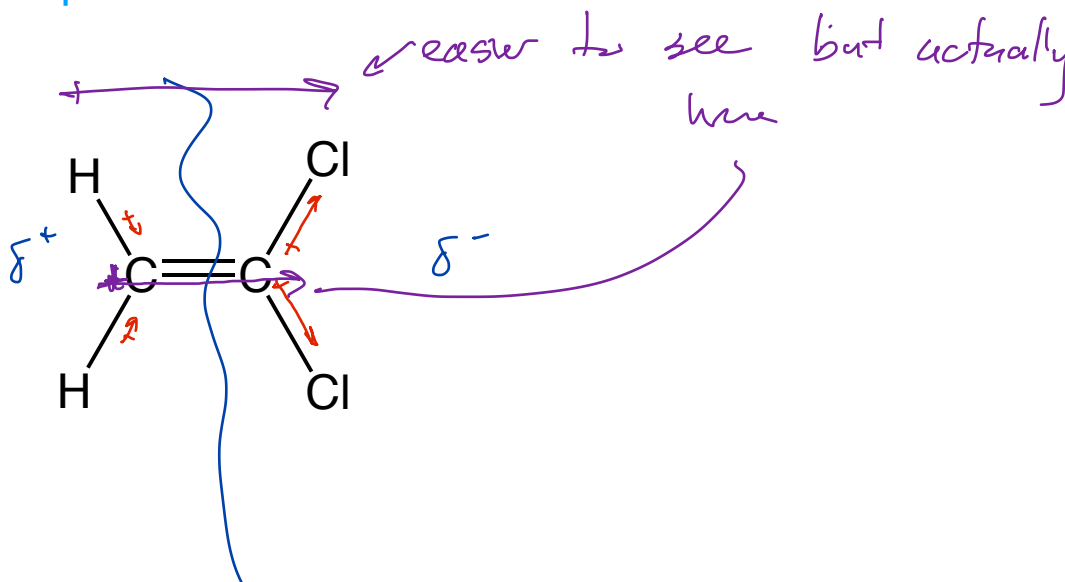
⁴ Ibid.

Draw Lewis Structure

Draw VSEPR Structure

Draw Dipole arrows on polar bonds

If (+) and (-) are on opposite sides (front and back, left and right, top and bottom, upper left and lower right) then the molecule is polar otherwise it isn't.

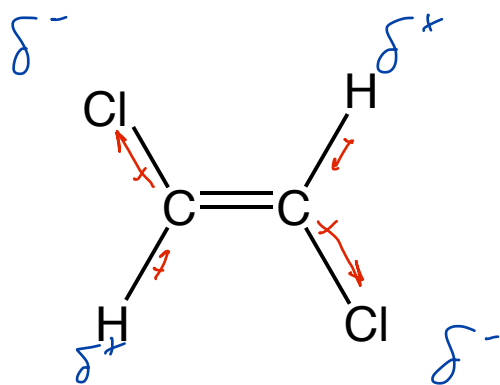


Draw Lewis Structure

Draw VSEPR Structure

Draw Dipole arrows on polar bonds

If (+) and (-) are on opposite sides (front and back, left and right, top and bottom, upper left and lower right) then the molecule is polar otherwise it isn't.



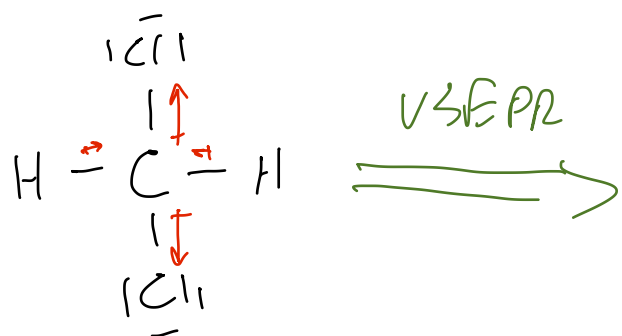
opposite charges
cannot be
separated by a
line/plane ...
non polar

Draw Lewis Structure

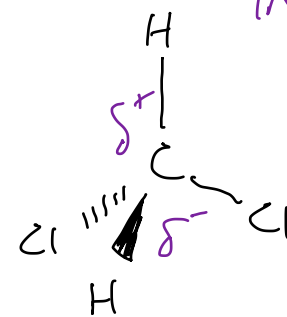
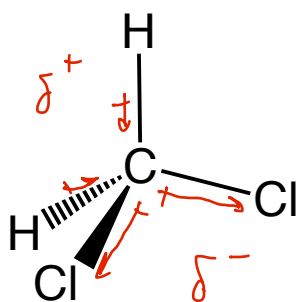
Draw VSEPR Structure

Draw Dipole arrows on polar bonds

If (+) and (-) are on opposite sides (front and back, left and right, top and bottom, upper left and lower right) then the molecule is polar otherwise it isn't.



VSEPR

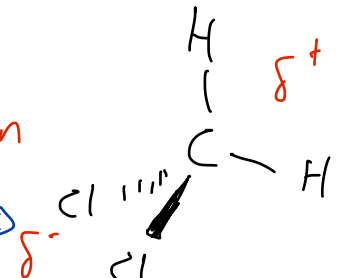


in this drawing
the δ^- is
behind the
screen

the δ^+ is
in front of
the screen

in front
of screen

group similar
atoms together
(if possible) to
visualize the charge separation
more clearly



Formal Charges

e⁻'s on atom before bonding

e⁻'s owned by atom after bonding

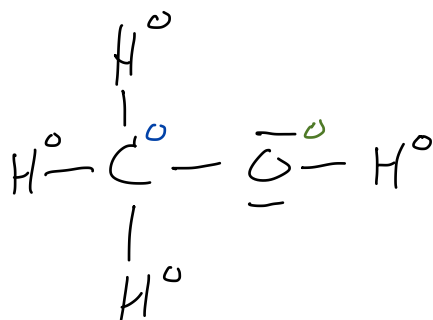
Formal Charge of an atom	=	# valence e ⁻ 's in the neutral atom	-	(# of lone pair electrons	+	1/2 of the e ⁻ 's in the bonds)
--------------------------	---	---	---	---	--------------------------	---	---	---



$$\begin{aligned}
 FC_H &= 1 - (0 + 1) \\
 &= 0
 \end{aligned}$$

$$\begin{aligned}
 FC_O &= 6 - (6 + 1) \\
 &= -1 \\
 &= 6 - 7 = -1
 \end{aligned}$$

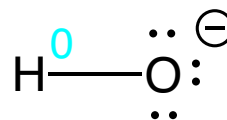
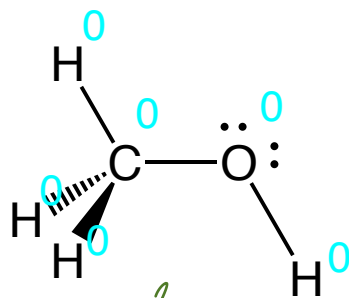
Formal Charge of an atom	=	# valence e ⁻ 's in the neutral atom	-	(# of lone pair electrons	+	1/2 of the e ⁻ 's in the bonds)
-----------------------------	---	--	---	---	-----------------------------	---	--	---

CH₃OH

$$FC_H = 1 - 1 = 0$$

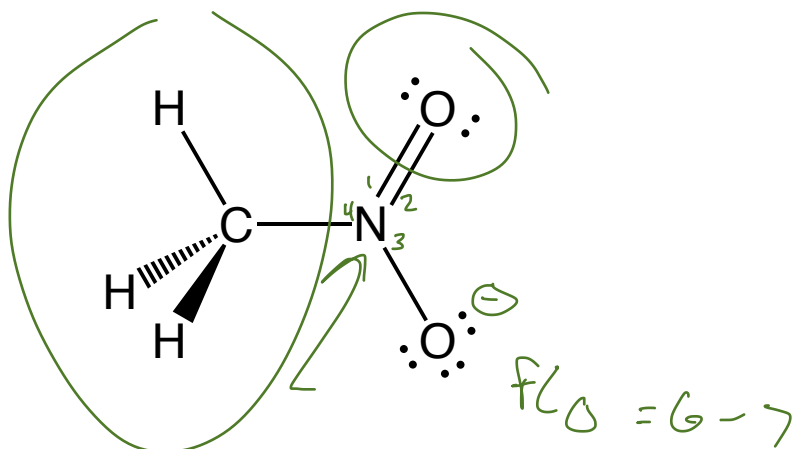
$$FC_C = 4 - (4) = 0$$

$$FC_O = 6 - 6 = 0$$



When you have the # of bonds you expect
based on an atom's position on the P. Table
 $FC = 0$

Formal Charge of an atom	=	# valence e ⁻ 's in the neutral atom	-	(# of lone pair electrons	+	1/2 of the e ⁻ 's in the bonds)
-----------------------------	---	--	---	---	-----------------------------	---	--	---



$$FC_N = 5 - (4) = 1$$