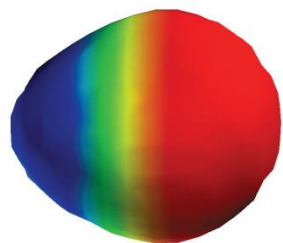


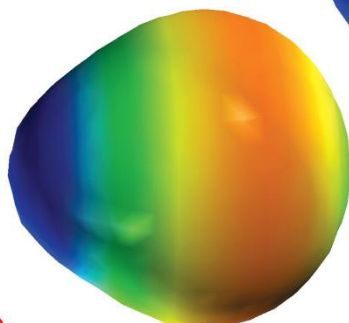
Organic Chemistry

2th Edition

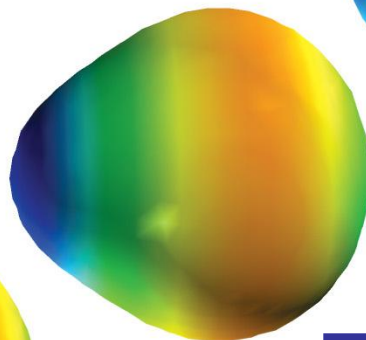
Paula Yurkanis Bruice



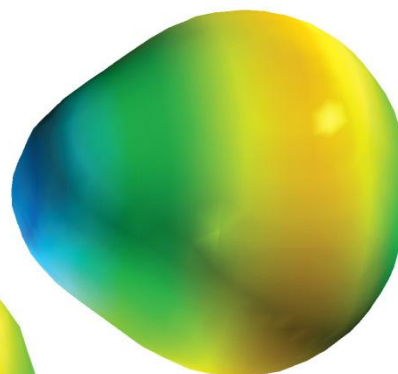
HF



HCl



HBr



HI

Chapter 1

Electronic Structure and Covalent Bonding

Irene Lee

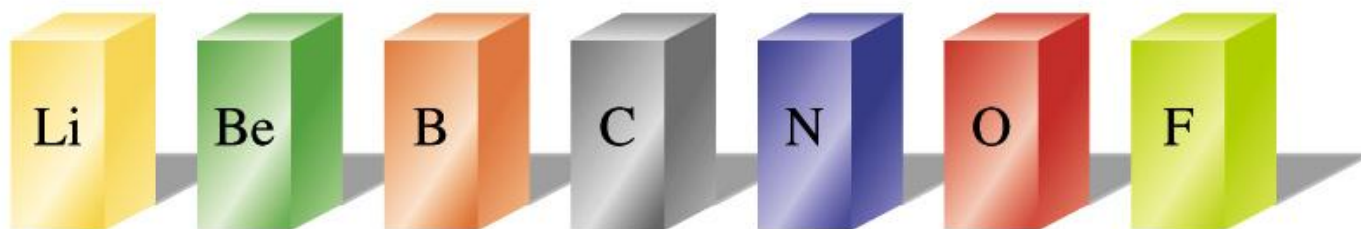
Case Western Reserve University

Cleveland, OH

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Organic Chemistry

- Organic compounds are compounds containing carbon
- Carbon neither readily gives up nor readily accepts electrons



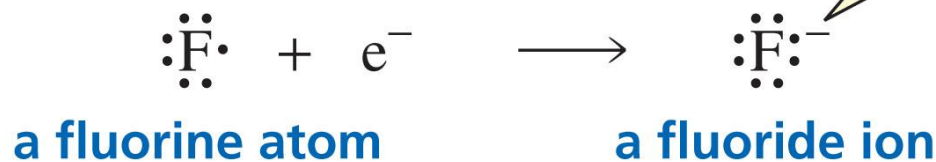
the second row of the periodic table

- Carbon shares electrons with other carbon atoms as well as with several different kinds of atoms

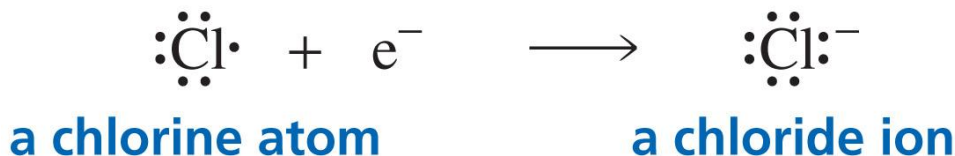
Lewis's theory: an atom will give up, accept, or share electrons in order to achieve a filled outer shell or an outer shell that contains eight electrons



lithium has lost an electron

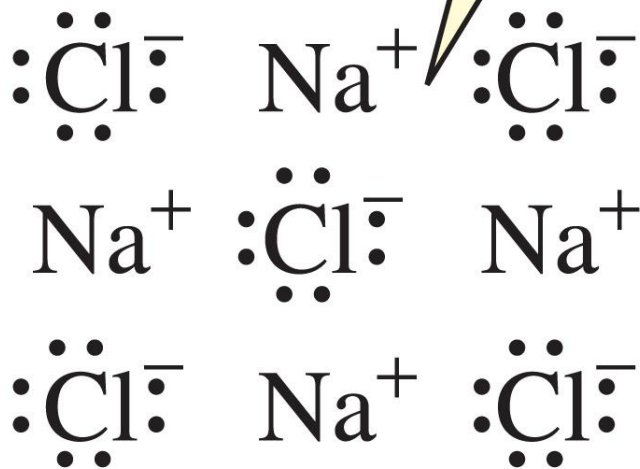


fluorine has gained an electron



Ionic Bonds Are Formed by the Transfer of Electrons

an ionic bond is the attraction between ions of opposite charges

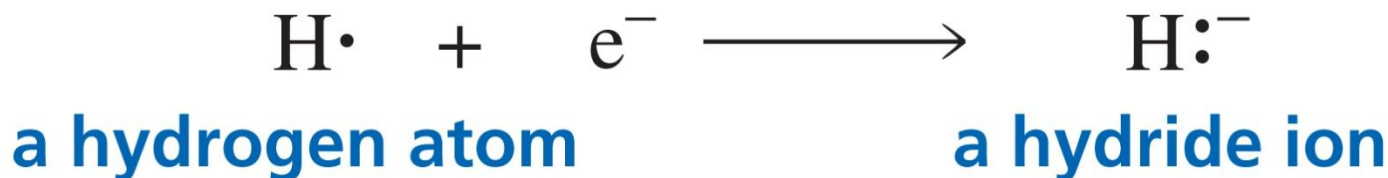
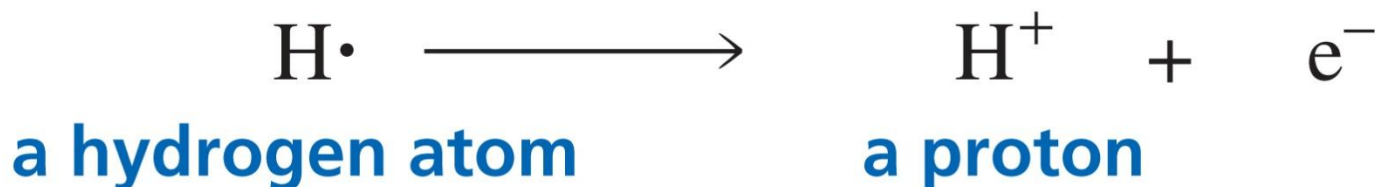
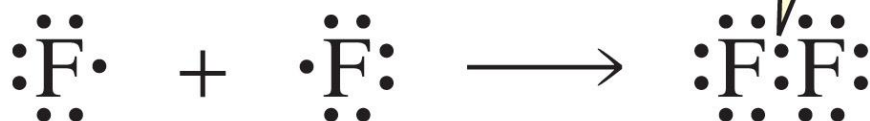


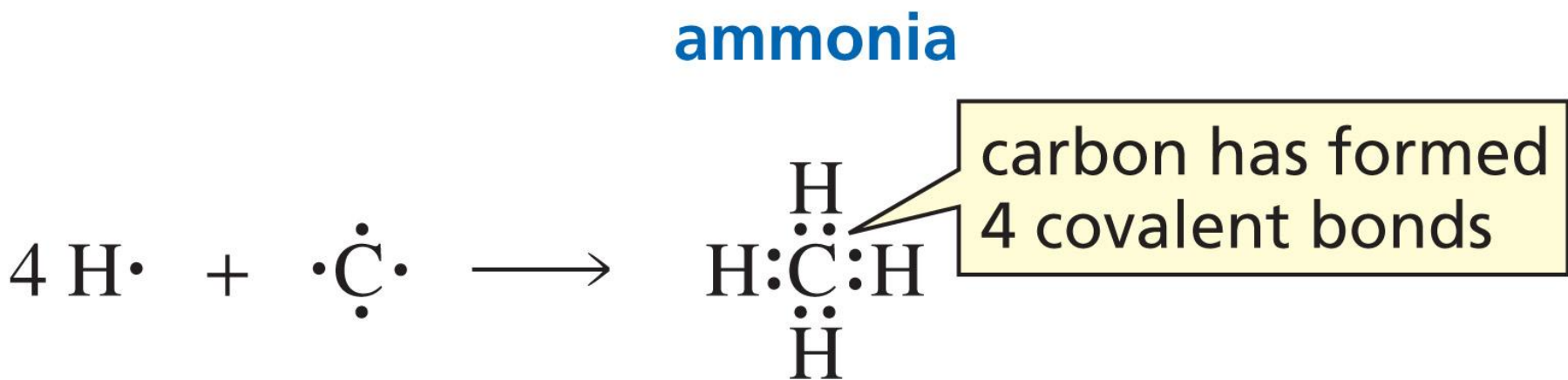
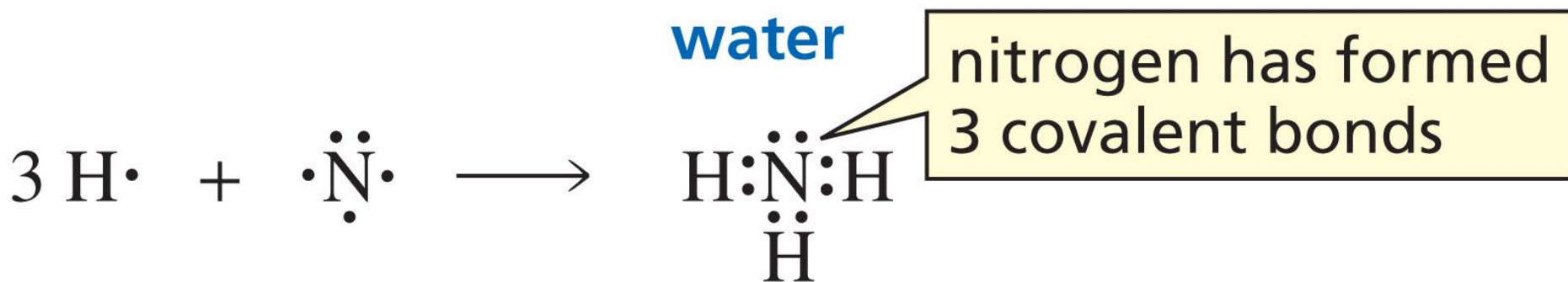
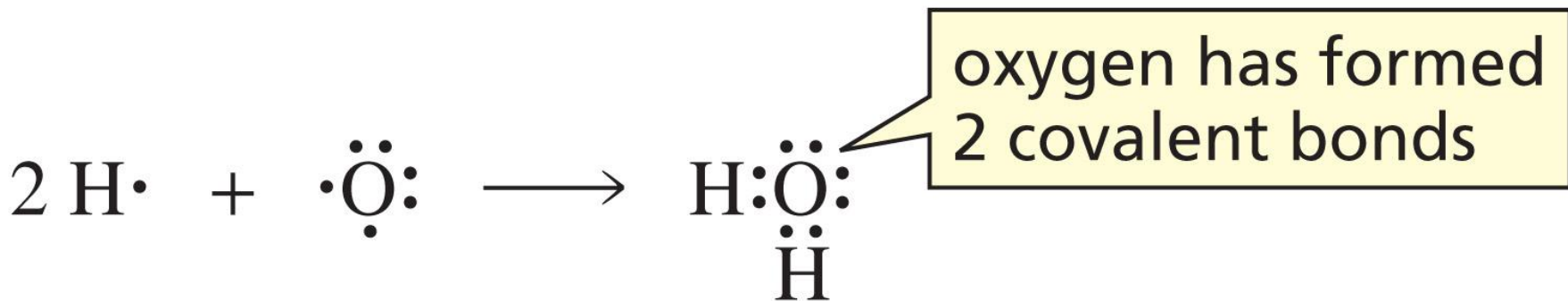
sodium chloride

Attractive forces between opposite charges are called electrostatic attractions

Covalent Bonds Are Formed by Sharing Electrons

a covalent bond is formed by sharing electrons



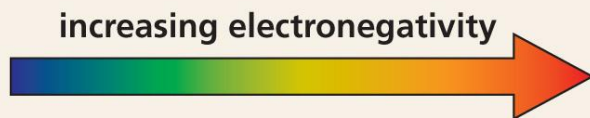


methane

- Equal sharing of electrons: nonpolar covalent bond (e.g., H₂)
- Sharing of electrons between atoms of different electronegativities: polar covalent bond (e.g., HF)

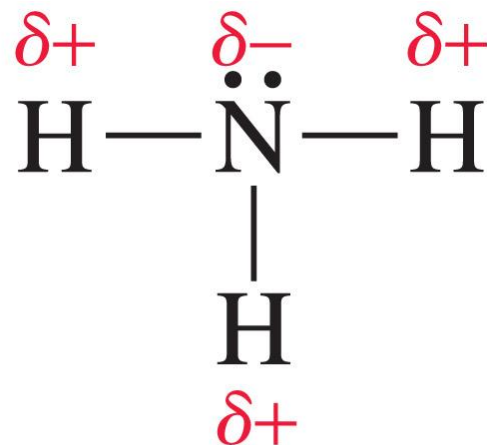
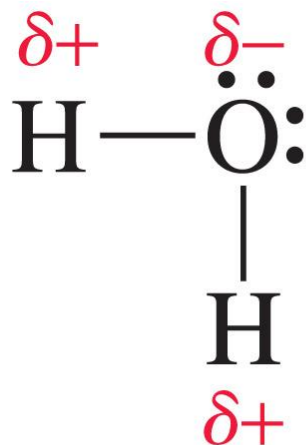
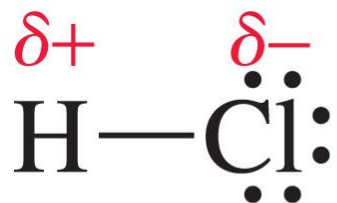
Table 1.3 The Electronegativities of Selected Elements^a

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



^aElectronegativity 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.

A polar covalent bond has a slight positive charge on one end and a slight negative charge on the other



the negative end
of the bond

continuum of bond types



ionic
bond



polar
covalent bond

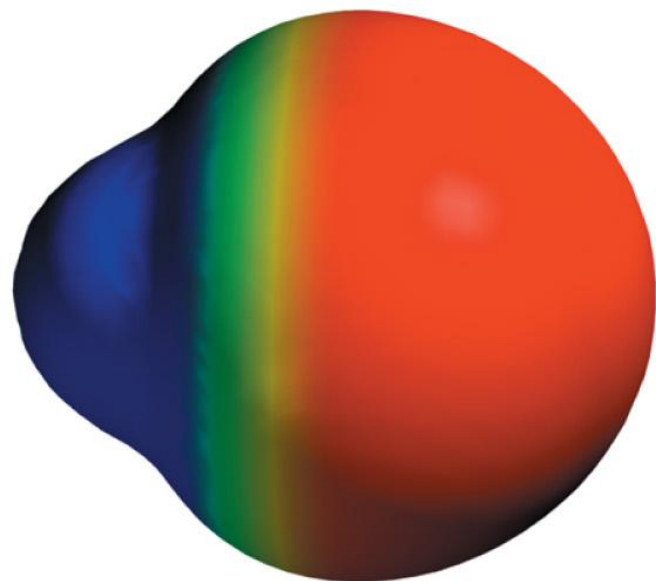


nonpolar
covalent bond

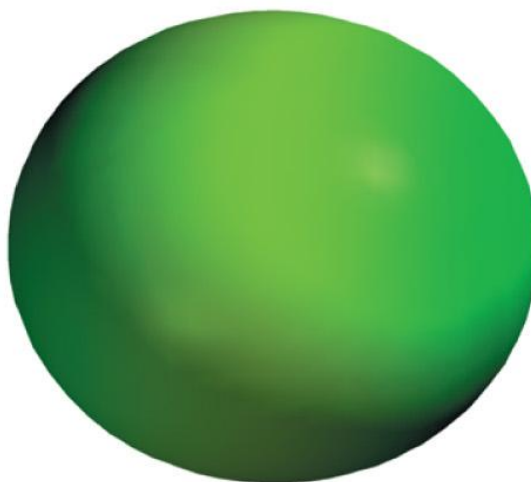


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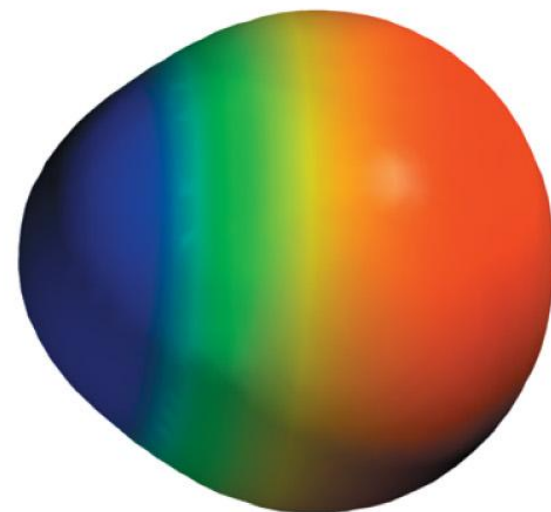
Electrostatic Potential Maps



LiH



H₂



HF

attracts positive charge



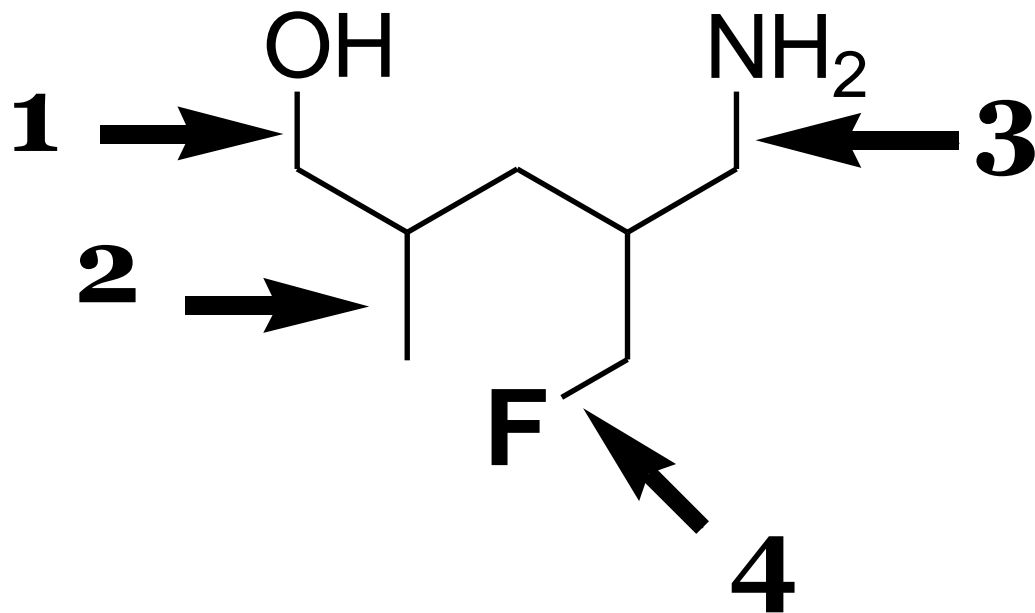
red • orange • yellow • green • blue

**most negative
electrostatic potential**

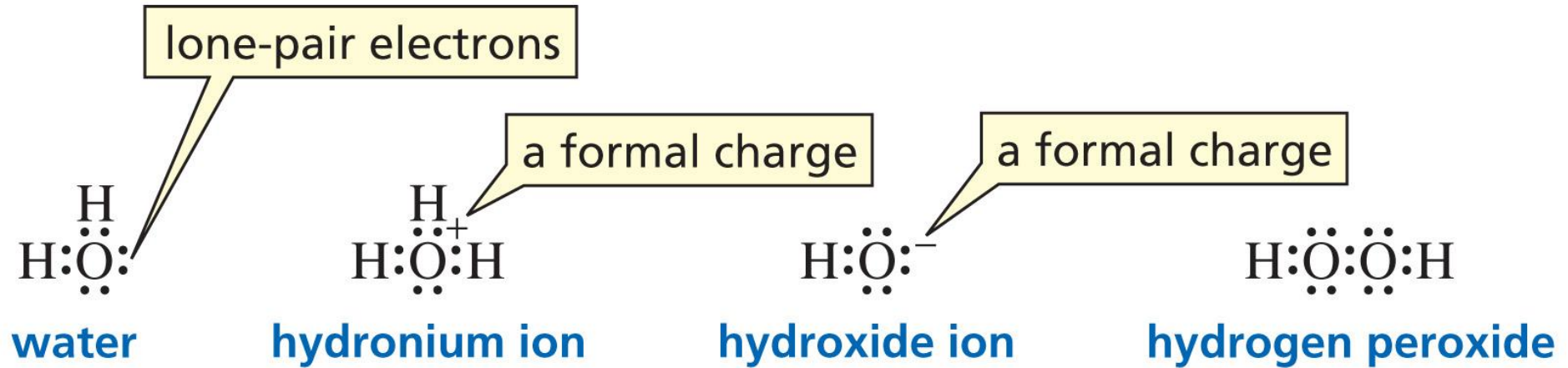
**most positive
electrostatic potential**

attracts negative charge

which bond is the most polar?



Lewis Structure



Formal charge =
number of valence electrons –
(number of lone pair electrons + 1/2 number of bonding electrons)

Important Bond Numbers

Neutral



one bond



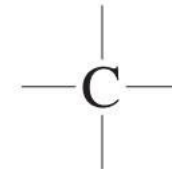
one bond



two bonds



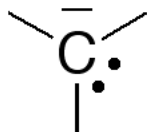
three bonds



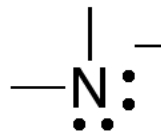
four bonds

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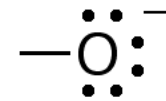
Anionic



three bonds



two bonds



one bond

Q Which of the following pairs will have the same bond number?

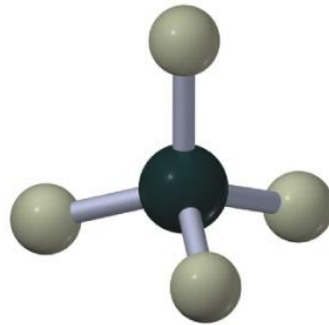
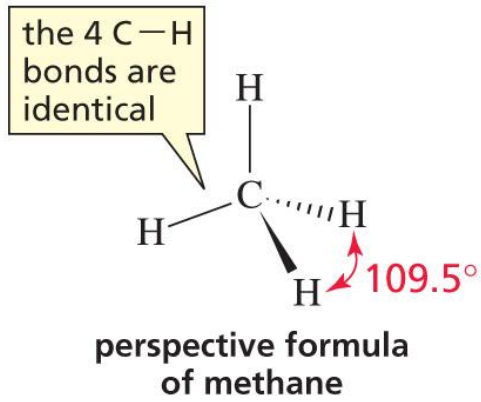
A) H and Cl

B) C and O

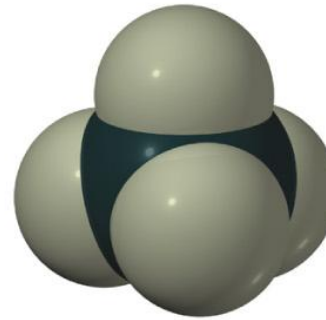
C) N^+ and C

D) B and F

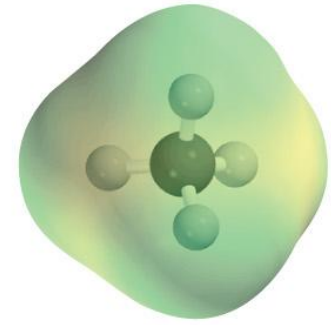
Bonding in Methane



ball-and-stick model of methane

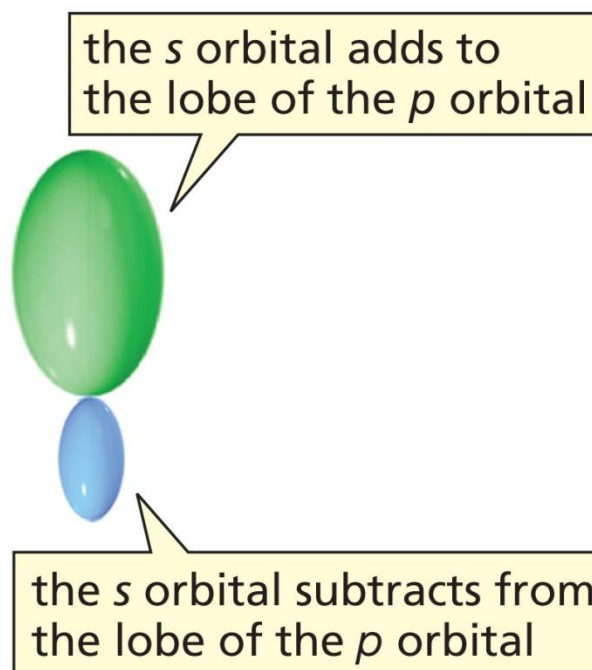
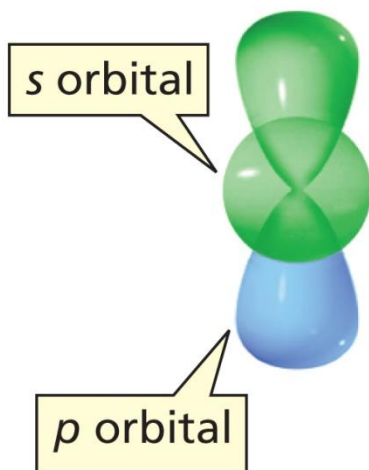
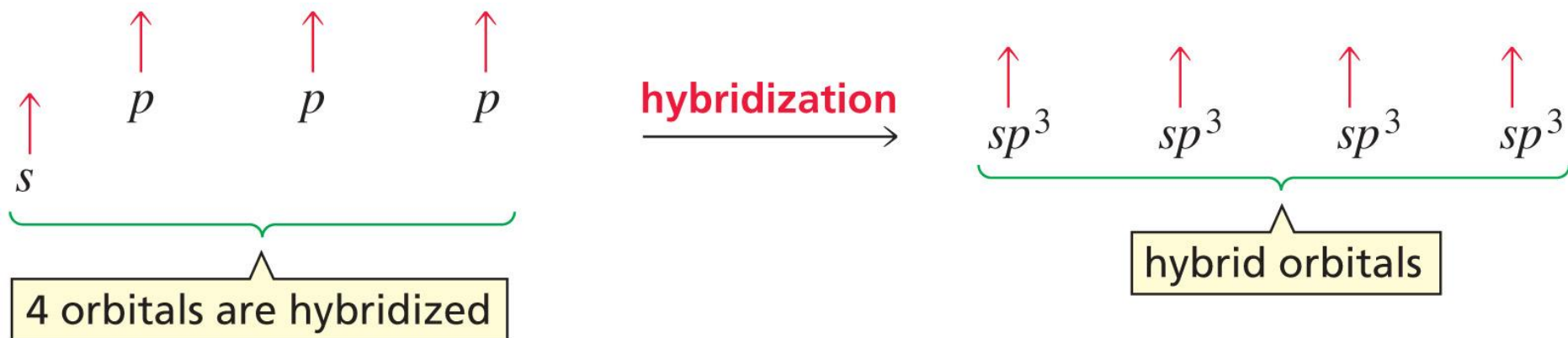


space-filling model of methane



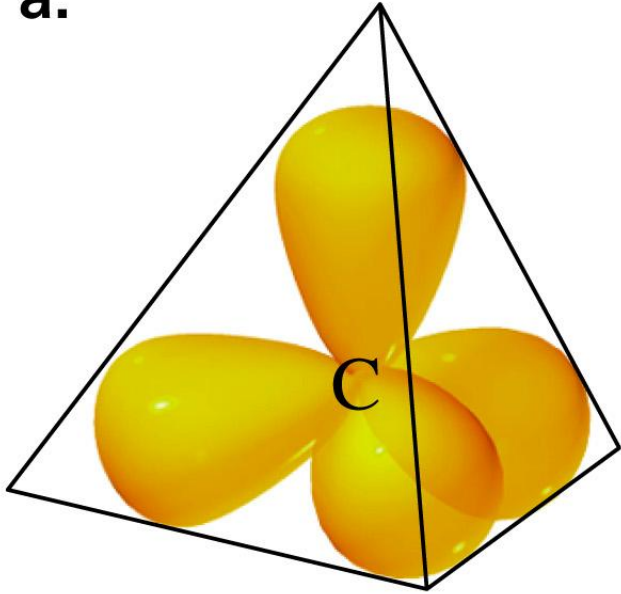
electrostatic potential map for methane

Hybridization of One s and Three p Orbitals

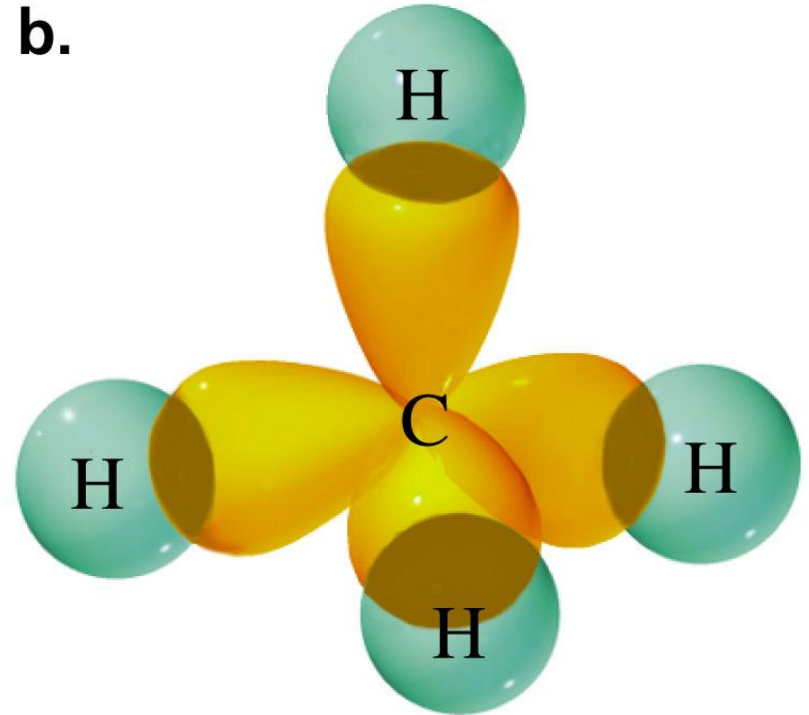


The orbitals used in bond formation determine the bond angles

a.

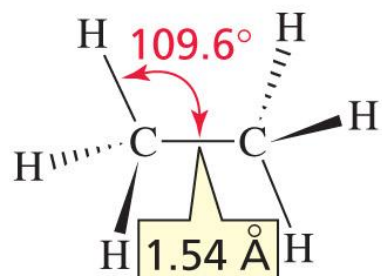


b.

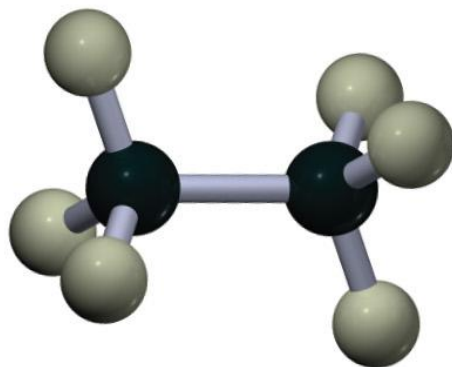


- Tetrahedral bond angle: 109.5°
- Electron pairs spread themselves into space as far from each other as possible

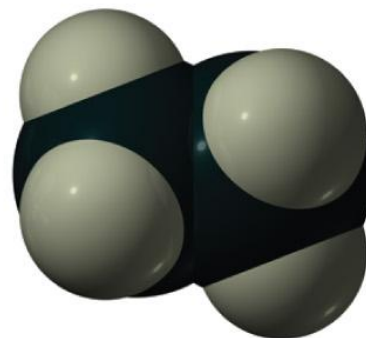
The Bonds in Ethane



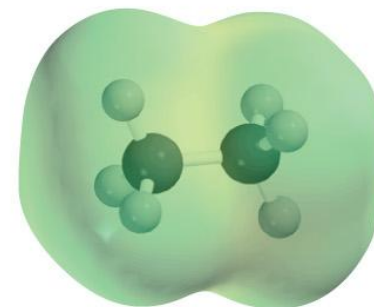
perspective formula of ethane



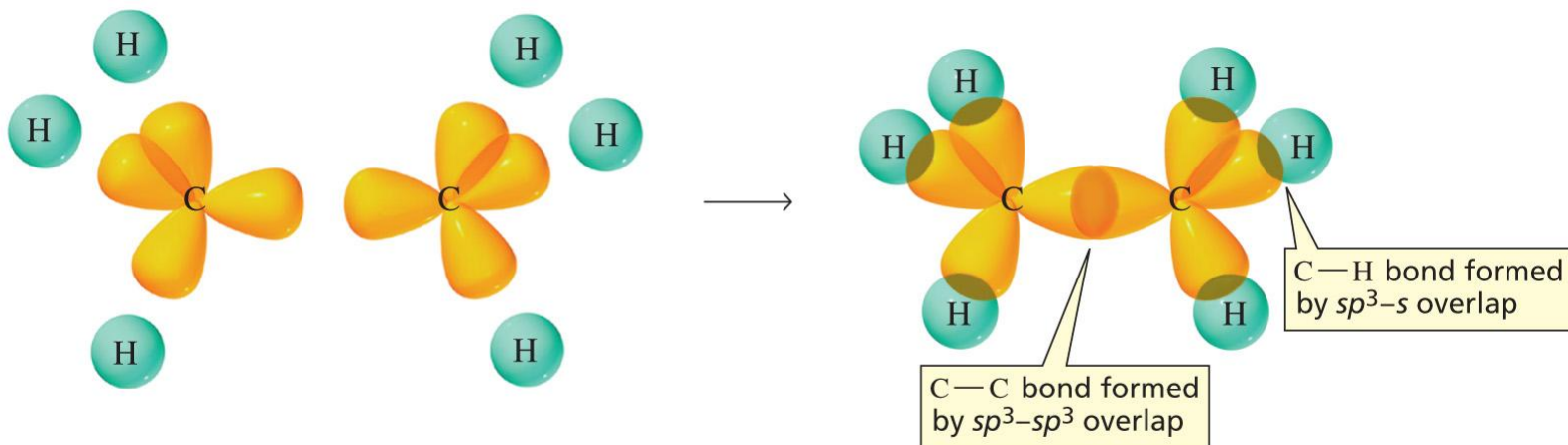
ball-and-stick model of ethane



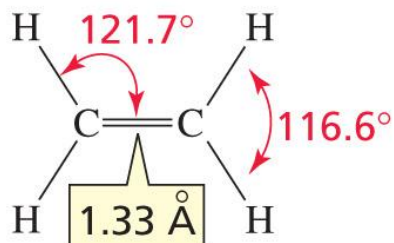
space-filling model of ethane



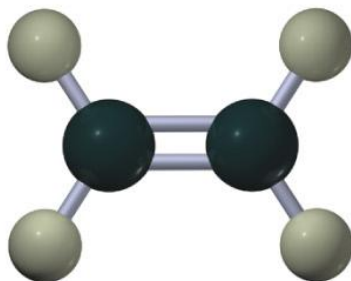
electrostatic potential map for ethane



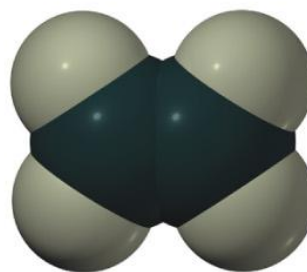
Bonding in Ethene: A Double Bond



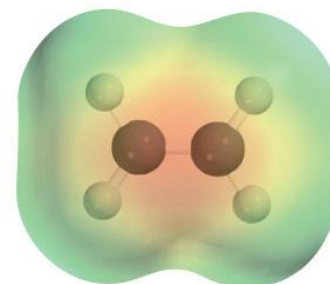
a double bond consists of one σ bond and one π bond



ball-and-stick model of ethene

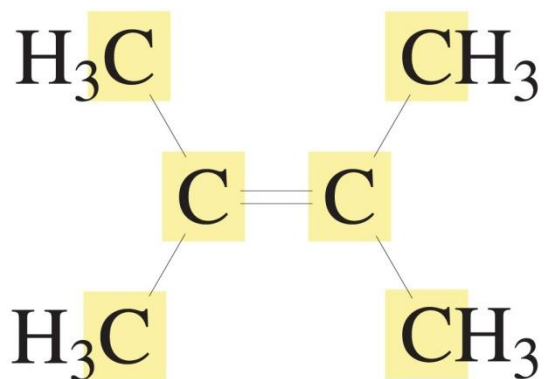


space-filling model of ethene



electrostatic potential map for ethene

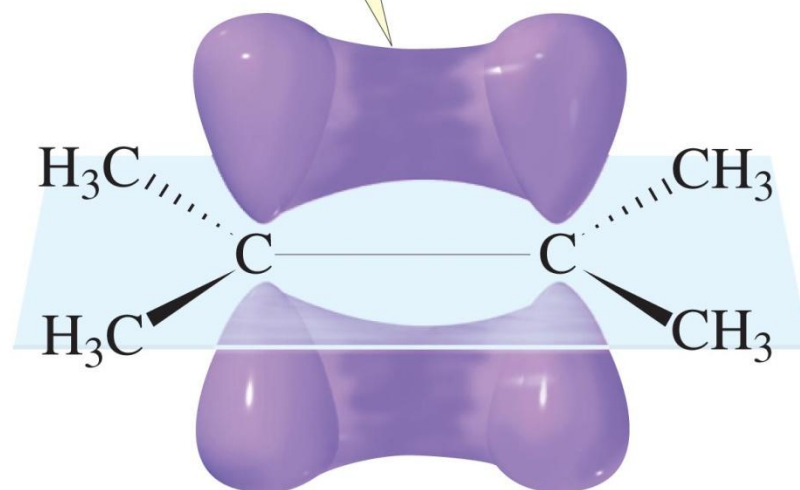
Structure of Alkene



**the six carbon atoms
are in the same plane**

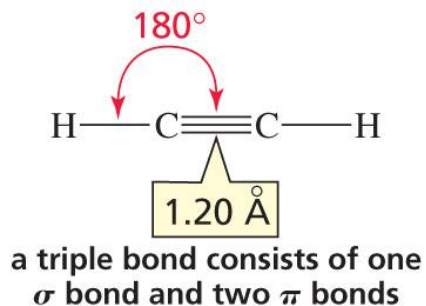
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p orbitals overlap to form
a π bond

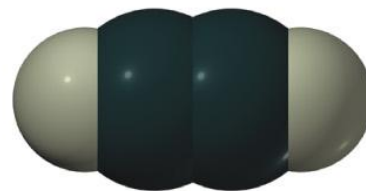


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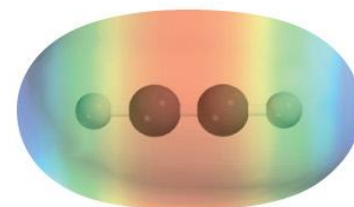
Bonding in Ethyne: A Triple Bond



ball-and-stick model of ethyne

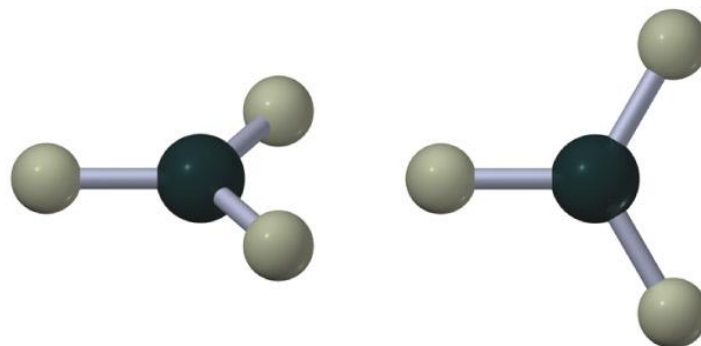
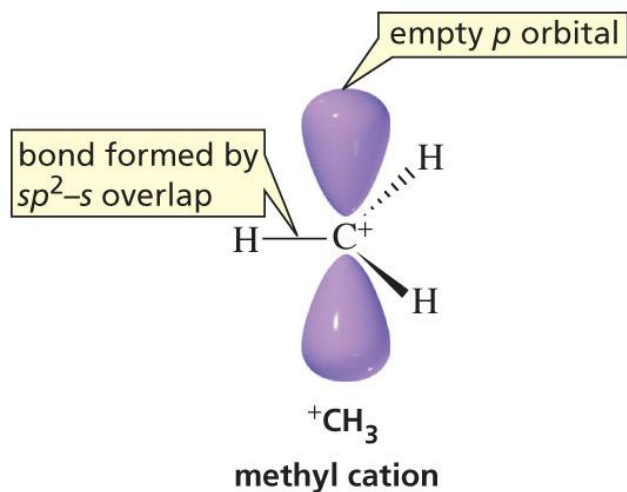


space-filling model of ethyne



electrostatic potential map for ethyne

Bonding in the Methyl Cation



angled side view

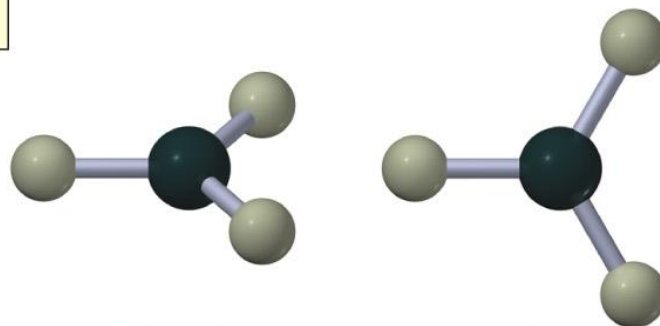
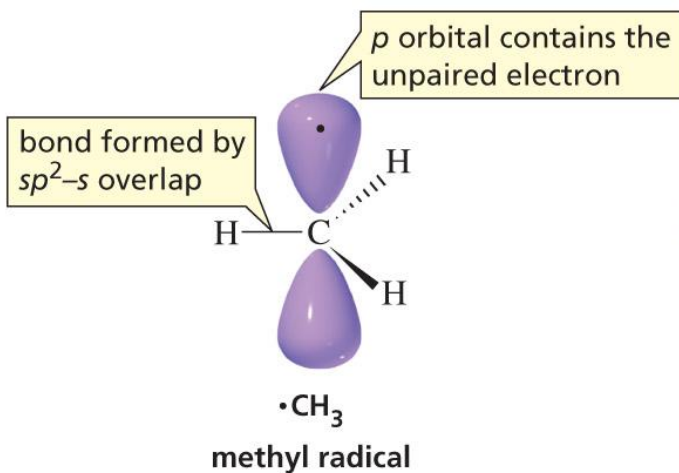
top view

ball-and-stick models of the methyl cation



electrostatic potential map
for the methyl cation

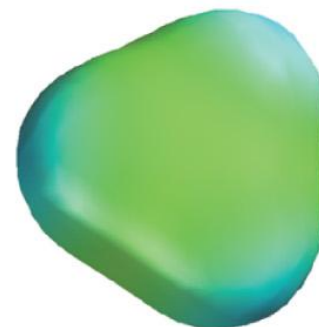
Bonding in the Methyl Radical



angled side view

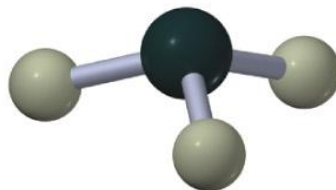
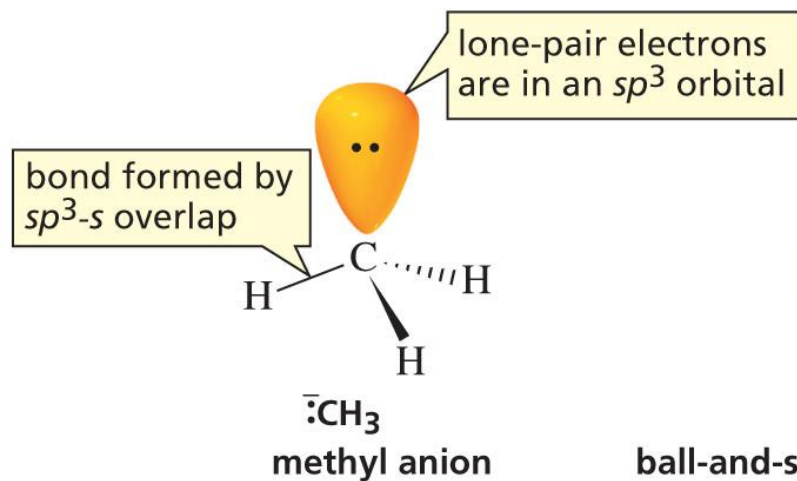
top view

ball-and-stick models of the methyl radical

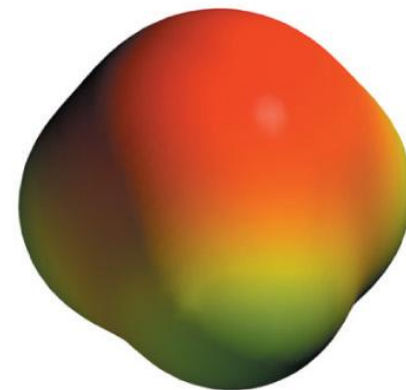


electrostatic potential map
for the methyl radical

Bonding in the Methyl Anion

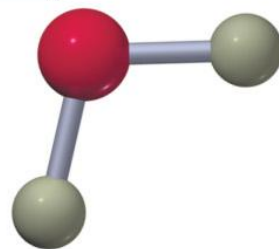
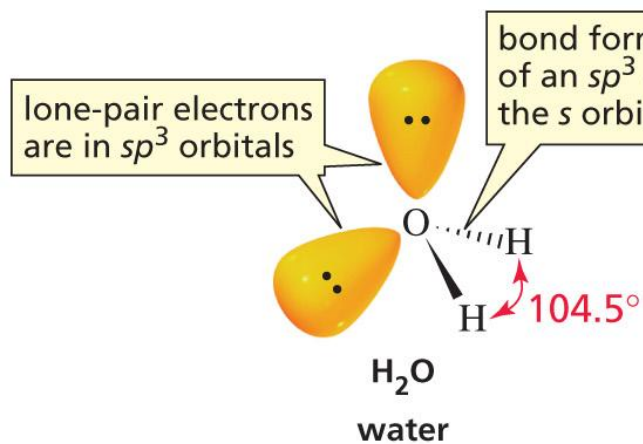


ball-and-stick model of the methyl anion

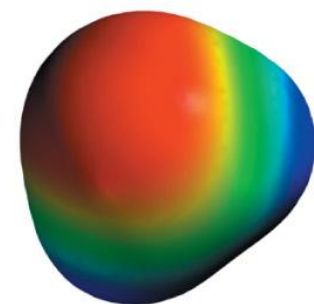


electrostatic potential map for the methyl anion

Bonding in Water



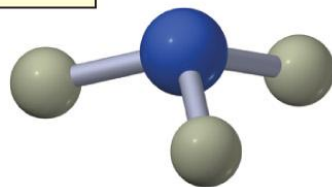
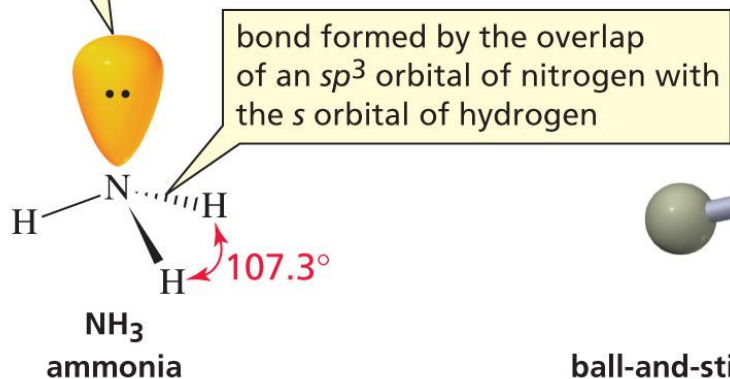
ball-and-stick model of water



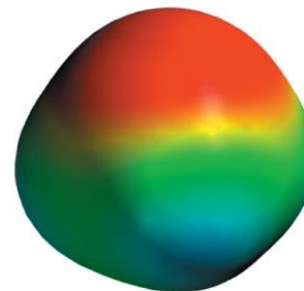
electrostatic potential map for water

Bonding in Ammonia and in the Ammonium Ion

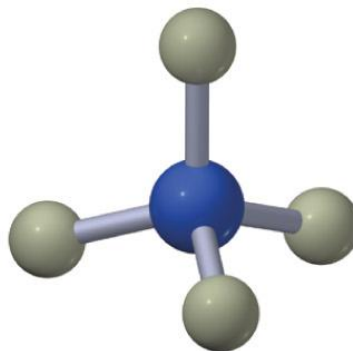
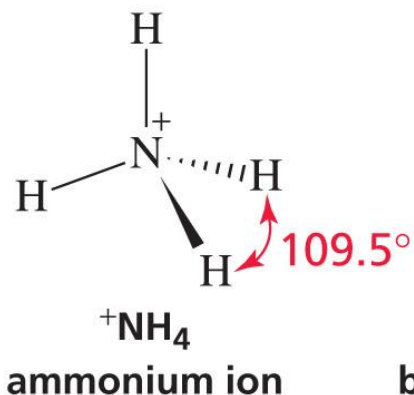
lone-pair electrons are in an sp^3 orbital



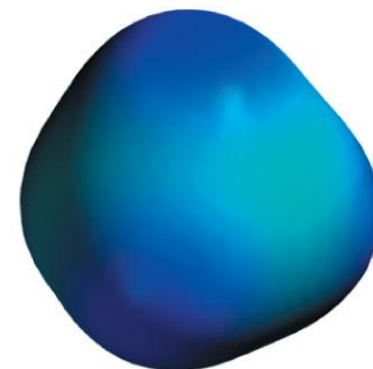
ball-and-stick model of ammonia



electrostatic potential map for ammonia

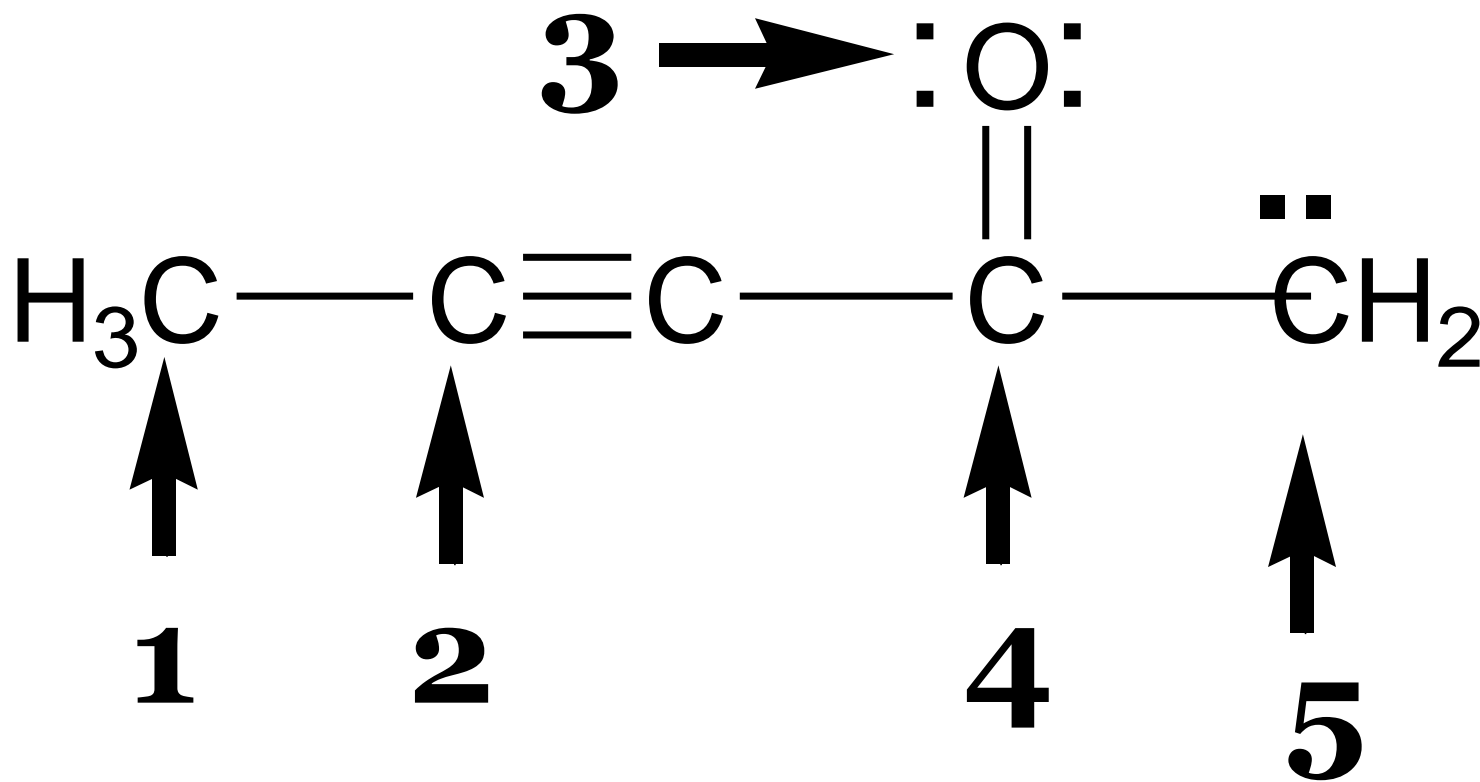


ball-and-stick model of the ammonium ion



electrostatic potential map for the ammonium ion

What is the hybridization of 1, 2, 3, 4, and 5?



Summary

- The shorter the bond, the stronger it is
- The greater the electron density in the region of orbital overlap, the stronger is the bond
- The more *s* character, the shorter and stronger is the bond
- The more *s* character, the larger is the bond angle