Universiti Malaysia PAHANG
Engineering • Technology • Creativity

Organic Chemistry

Chemical Bonding and Structure

by

Dr. Seema Zareen & Dr. Izan Izwan Misnon
Faculty Industrial Science & Technology
seema@ump.edu.my & iezwan@ump.edu.my

Expected Outcomes

In the end of this chapter, student will have the ability to:

- Draw Lewis structure
- Identify ionic and covalent bond in a compounds
- Differentiate isomers and resonance Lewis drawing
- Explain characteristic and properties of constitutional isomers, enantiomers, diastereoisomers, and racemic mixture
- Predict the shape of molecules
- Draw condensed structures and skeletal structure

Contents

- Bonding
- Lewis Structure
- Resonance
- Stereochemistry
- Molecular shapes
- Drawing organic structure

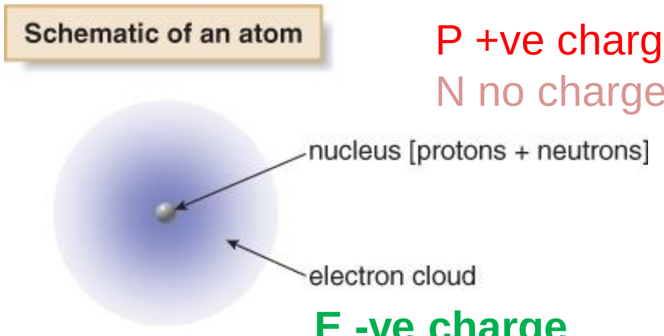
Organic Chemistry

- The branch of Science in which we study of the carbon containing compounds is called Organic Chemistry.

OR

- The study of hydrocarbons and their derivatives known as Organic Chemistry.
- Over 10 million structures have been identified
  - about 1000 new ones are identified each day! (Scifinder website)
- C is a small atom
  - it forms single, double, and triple bonds
  - it is intermediate in electronegativity (2.5)
  - it forms strong bonds with C, H, O, N, and some metals

Structure and Bonding



P +ve charge

N no charge

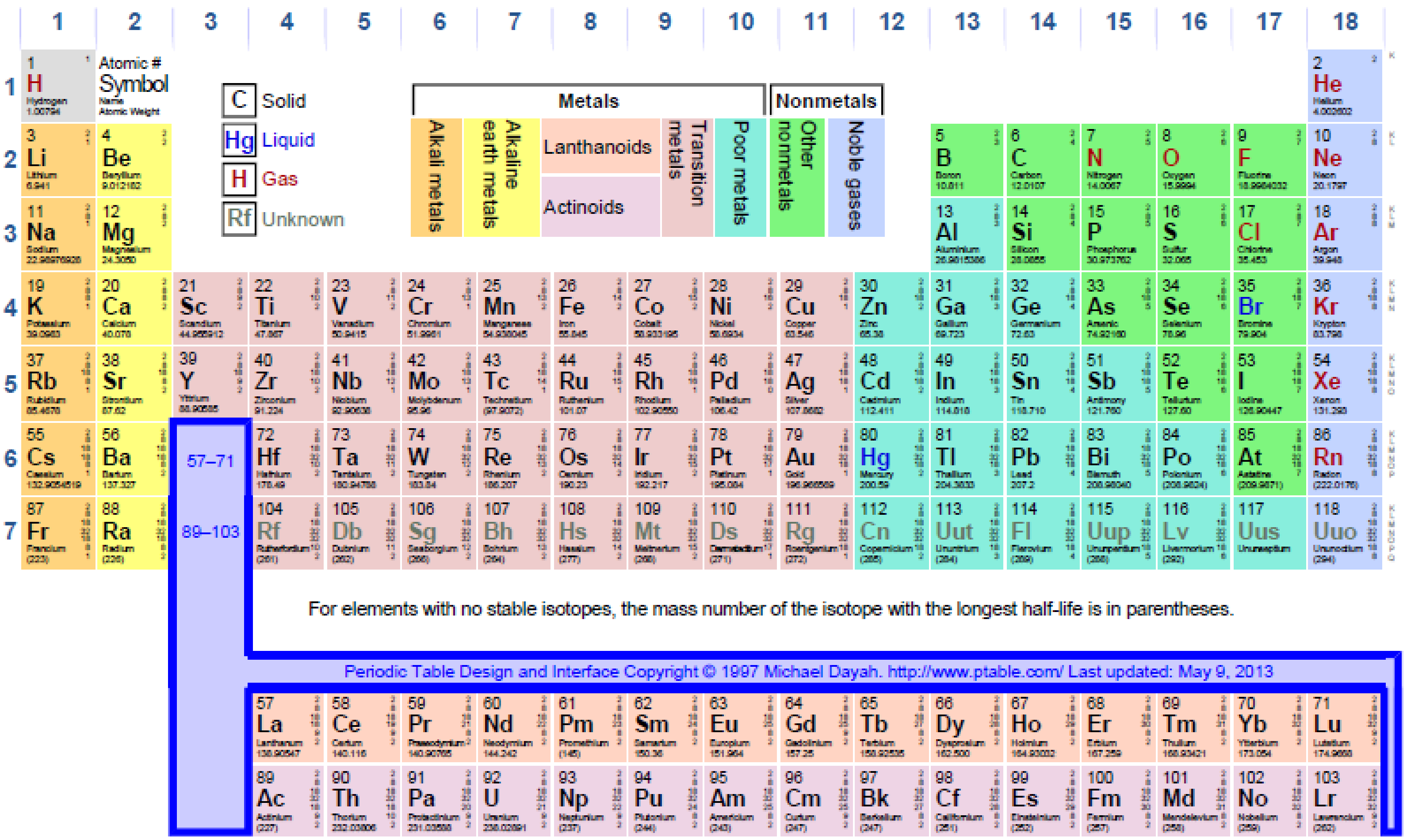
E -ve charge

- **The nucleus contains positively charged protons and uncharged neutrons.**
- **The electron cloud is composed of negatively charged electrons.**

N = A – Z

∴ N = No. of proton
A = Atomic number = No. of proton = No. of electron
Z = Mass Number or Atomic Mass = No. of proton + No. of neutron

Periodic Table of Elements



Atomic #
Symbol
Name
Atomic Weight

C Solid
Hg Liquid
H Gas
Rf Unknown

Metals: Alkali metals, Alkaline earth metals, Lanthanoids, Actinoids, Transition metals, Poor metals
Nonmetals: Other nonmetals, Noble gases

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H Hydrogen																	2 He Helium
2	3 Li Lithium	4 Be Beryllium											5 B Boron	6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne Neon
3	11 Na Sodium	12 Mg Magnesium											13 Al Aluminium	14 Si Silicon	15 P Phosphorus	16 S Sulfur	17 Cl Chlorine	18 Ar Argon
4	19 K Potassium	20 Ca Calcium	21 Sc Scandium	22 Ti Titanium	23 V Vanadium	24 Cr Chromium	25 Mn Manganese	26 Fe Iron	27 Co Cobalt	28 Ni Nickel	29 Cu Copper	30 Zn Zinc	31 Ga Gallium	32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton
5	37 Rb Rubidium	38 Sr Strontium	39 Y Yttrium	40 Zr Zirconium	41 Nb Niobium	42 Mo Molybdenum	43 Tc Technetium	44 Ru Ruthenium	45 Rh Rhodium	46 Pd Palladium	47 Ag Silver	48 Cd Cadmium	49 In Indium	50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe Xenon
6	55 Cs Caesium	56 Ba Barium	57–71	72 Hf Hafnium	73 Ta Tantalum	74 W Tungsten	75 Re Rhenium	76 Os Osmium	77 Ir Iridium	78 Pt Platinum	79 Au Gold	80 Hg Mercury	81 Tl Thallium	82 Pb Lead	83 Bi Bismuth	84 Po Polonium	85 At Astatine	86 Rn Radon
7	87 Fr Francium	88 Ra Radium	89–103	104 Rf Rutherfordium	105 Db Dubnium	106 Sg Seaborgium	107 Bh Bohrium	108 Hs Hassium	109 Mt Meitnerium	110 Ds Darmstadtium	111 Rg Roentgenium	112 Cn Copernicium	113 Uut Ununtrium	114 Fl Flerovium	115 Uup Ununpentium	116 Lv Livermorium	117 Uus Ununseptium	118 Uuo Ununoctium

For elements with no stable isotopes, the mass number of the isotope with the longest half-life is in parentheses.

57 La Lanthanum	58 Ce Cerium	59 Pr Praseodymium	60 Nd Neodymium	61 Pm Promethium	62 Sm Samarium	63 Eu Europium	64 Gd Gadolinium	65 Tb Terbium	66 Dy Dysprosium	67 Ho Holmium	68 Er Erbium	69 Tm Thulium	70 Yb Ytterbium	71 Lu Lutetium
89 Ac Actinium	90 Th Thorium	91 Pa Protactinium	92 U Uranium	93 Np Neptunium	94 Pu Plutonium	95 Am Americium	96 Cm Curium	97 Bk Berkelium	98 Cf Californium	99 Es Einsteinium	100 Fm Fermium	101 Md Mendelevium	102 No Nobelium	103 Lr Lawrencium

Michael Dayah

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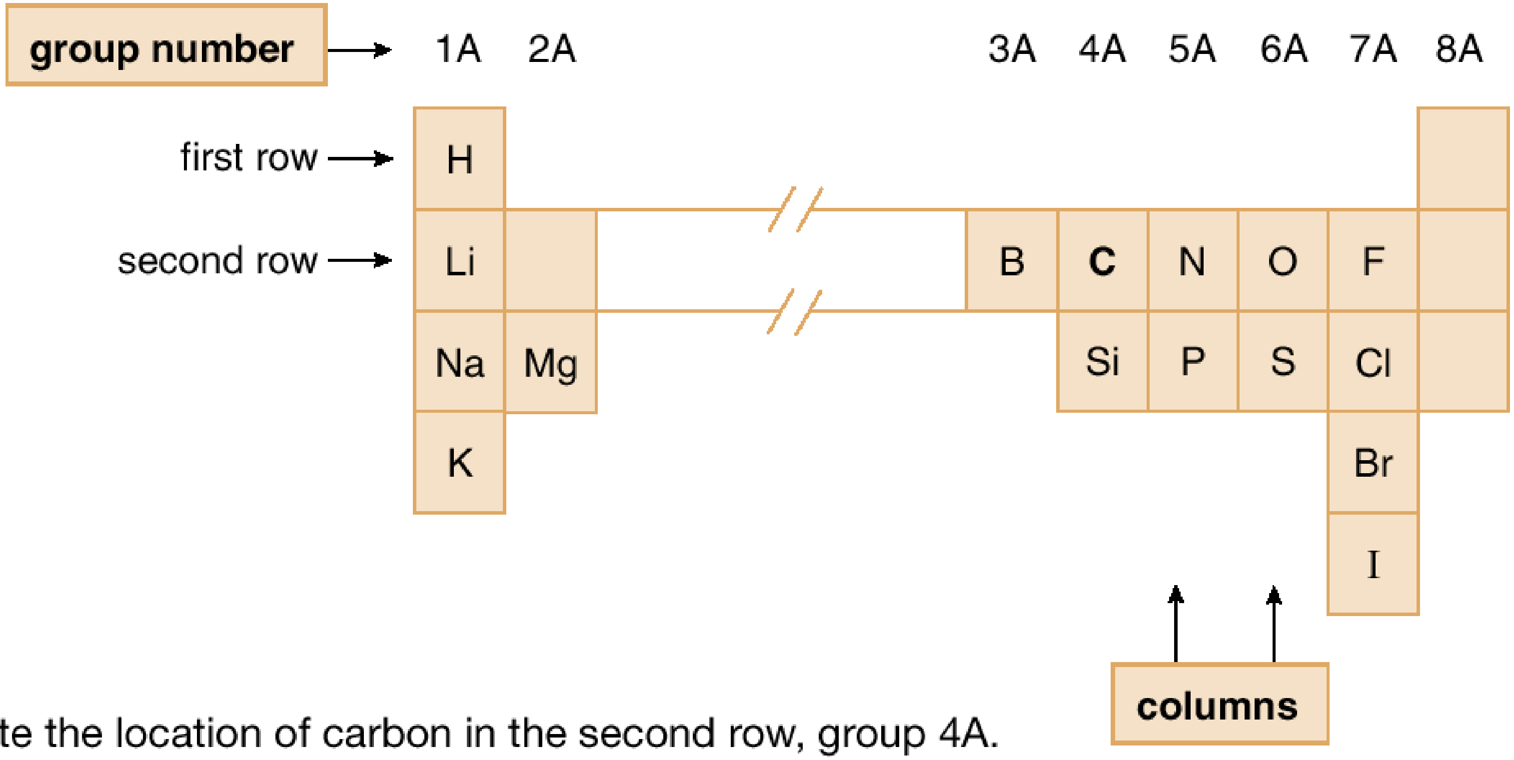
michael@dayah.com

Structure and Bonding

The Periodic Table

A periodic table of the common elements seen in organic chemistry

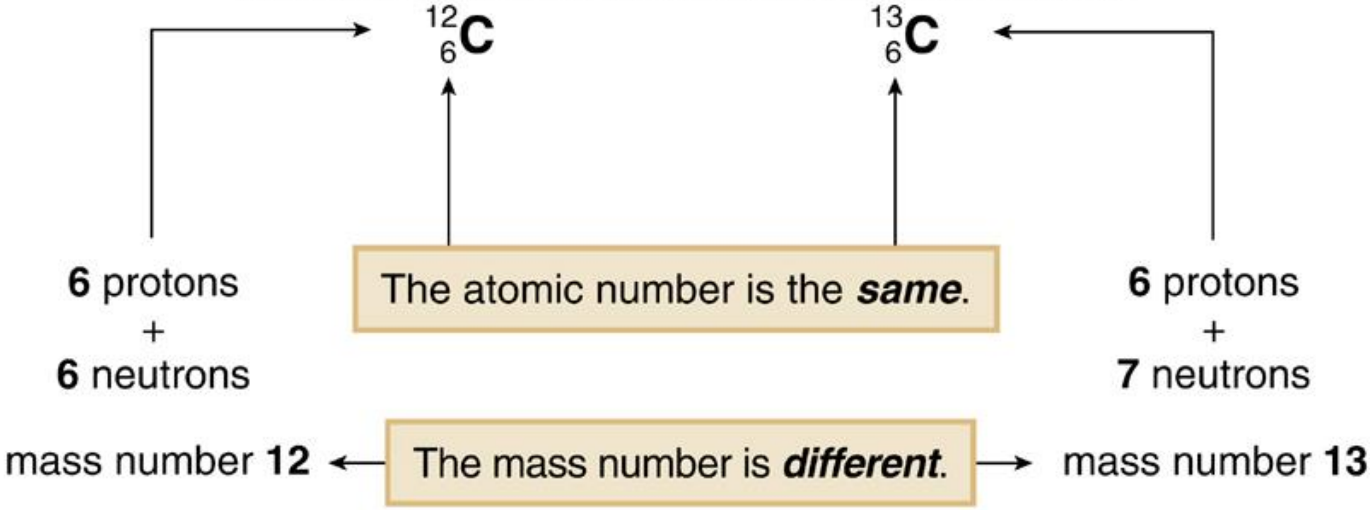


- Note the location of carbon in the second row, group 4A.

Structure and Bonding

The Periodic Table

A comparison of two isotopes of the element carbon



$^{12}_{6}C$

$^{13}_{6}C$

6 protons
+
6 neutrons

The atomic number is the ***same***.

6 protons
+
7 neutrons

mass number **12** ← The mass number is ***different***. → mass number **13**

- Elements in the same row are similar in size.
- Elements in the same column have similar electronic and chemical properties.

Structure and Bonding

Second Row Elements

- Since each of the four orbitals available in the second shell can hold two electrons, there is a maximum capacity of eight electrons for elements in the second row.

- The second row of the periodic chart consists of eight elements, obtained by adding electrons to the 2*s* and three 2*p* orbitals.

group number	1A	2A	3A	4A	5A	6A	7A	8A
second row	Li	Be	B	C	N	O	F	Ne
number of valence electrons	1	2	3	4	5	6	7	8

Structure and Bonding

- **Chemical bond:** attractive force holding two or more atoms together.

There are 2 extreme forms of connecting or bonding atoms:

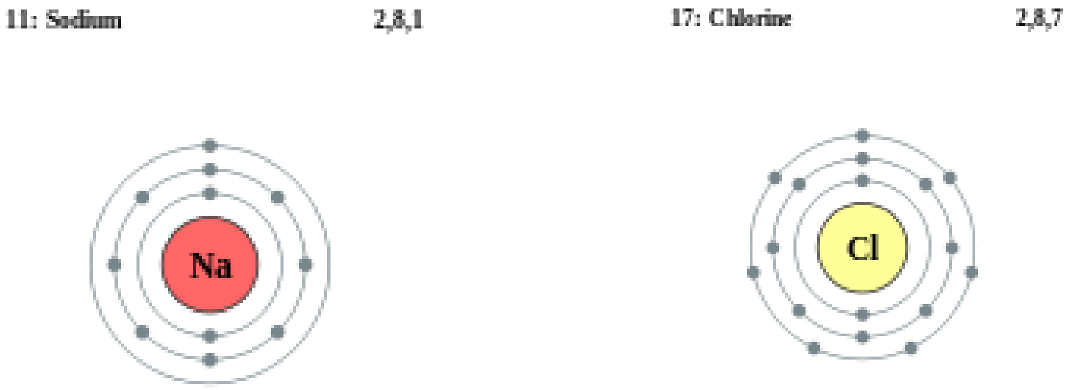
- **Ionic Bond**—complete transfer of electrons from one atom to another
- **Covalent Bond**—electrons shared between atoms

Most bonds are somewhere in between.

Structure and Bonding

IONIC BOND FORMATION

Sodium loss one electron get positive charge and Chlorine get electron be come negative charge, both ions form ionic bond through electrostatic force of attraction



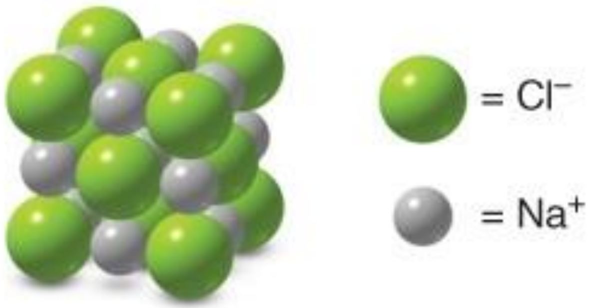
2,8,1 + 2,8,7 → 2,8,8

$Na^+ + Cl^- \longrightarrow NaCl$

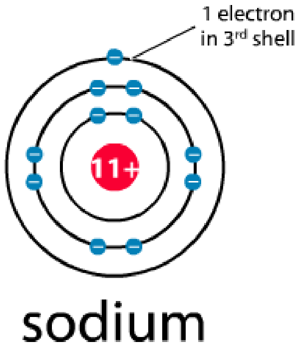
Structure and Bonding

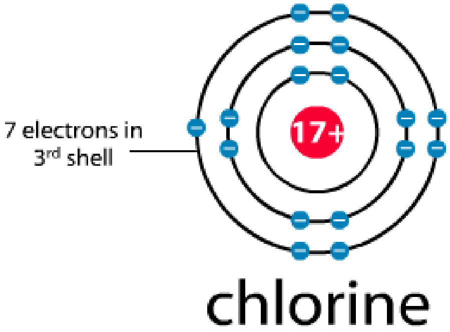
- **An ionic bond generally occurs when elements on the far left side of the periodic table combine with elements on the far right side, ignoring noble gases.**
- **A positively charged cation formed from the element on the left side attracts a negatively charged anion formed from the element on the right side. An example is sodium chloride, NaCl.**

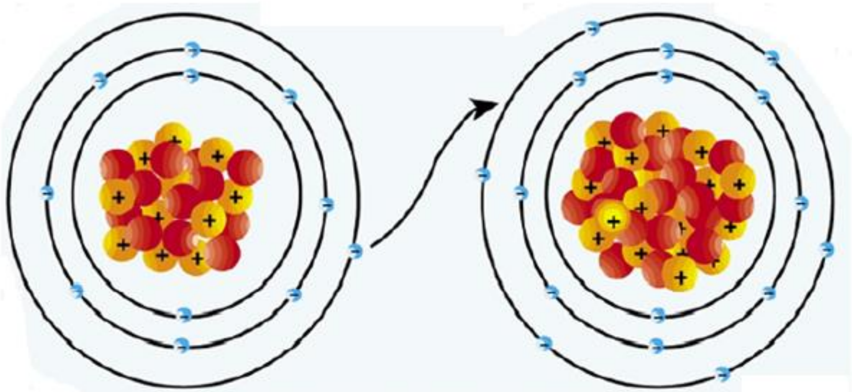
NaCl—An ionic crystalline lattice



Structure and Bonding







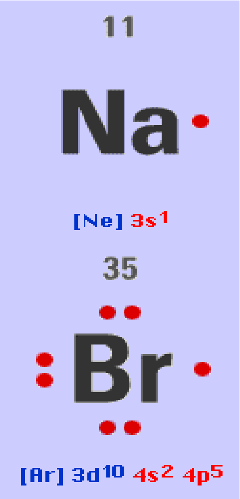
Ionic Bond – Sodium Takes an Electron
From Chlorine to Make a Salt Molecule

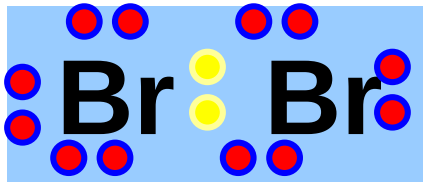
Covalent Bonding

Covalent bond is the sharing of the VALENCE ELECTRONS of each atom in a bond

Recall: Electrons are divided between core and valence electrons.

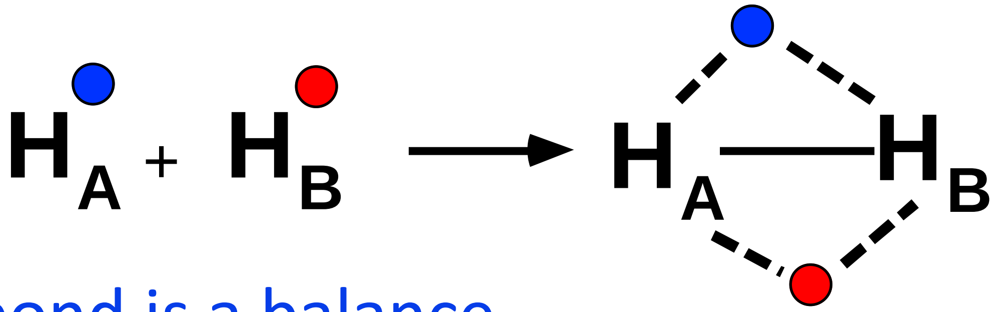
ATOM		core	valence
Na	$1s^2\ 2s^2\ 2p^6\ 3s^1$	[Ne]	$3s^1$
Br	$[Ar]\ 3d^{10}\ 4s^2\ 4p^5$	$[Ar]\ 3d^{10}$	$4s^2\ 4p^5$



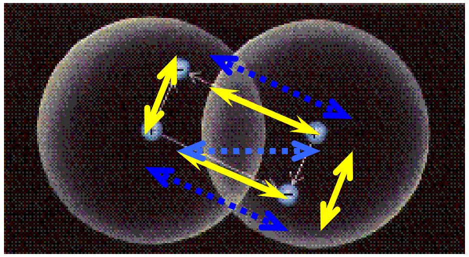


Covalent Bonding

The bond arises from the mutual attraction of 2 nuclei for the same electrons.



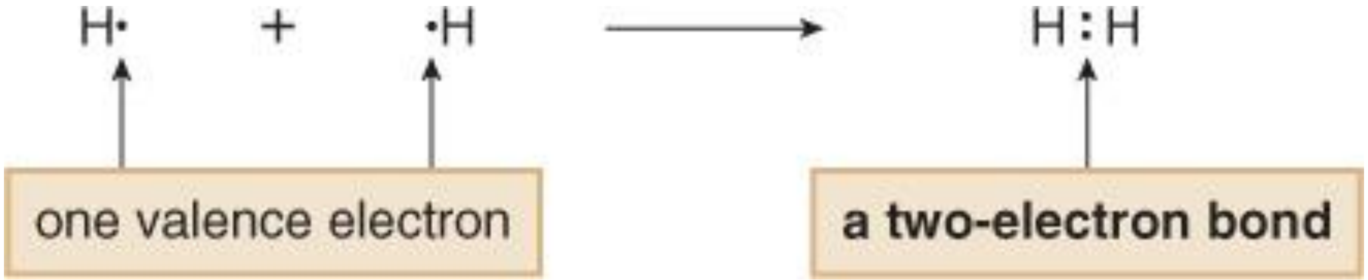
A covalent bond is a balance of attractive and repulsive forces.



Covalent Bond

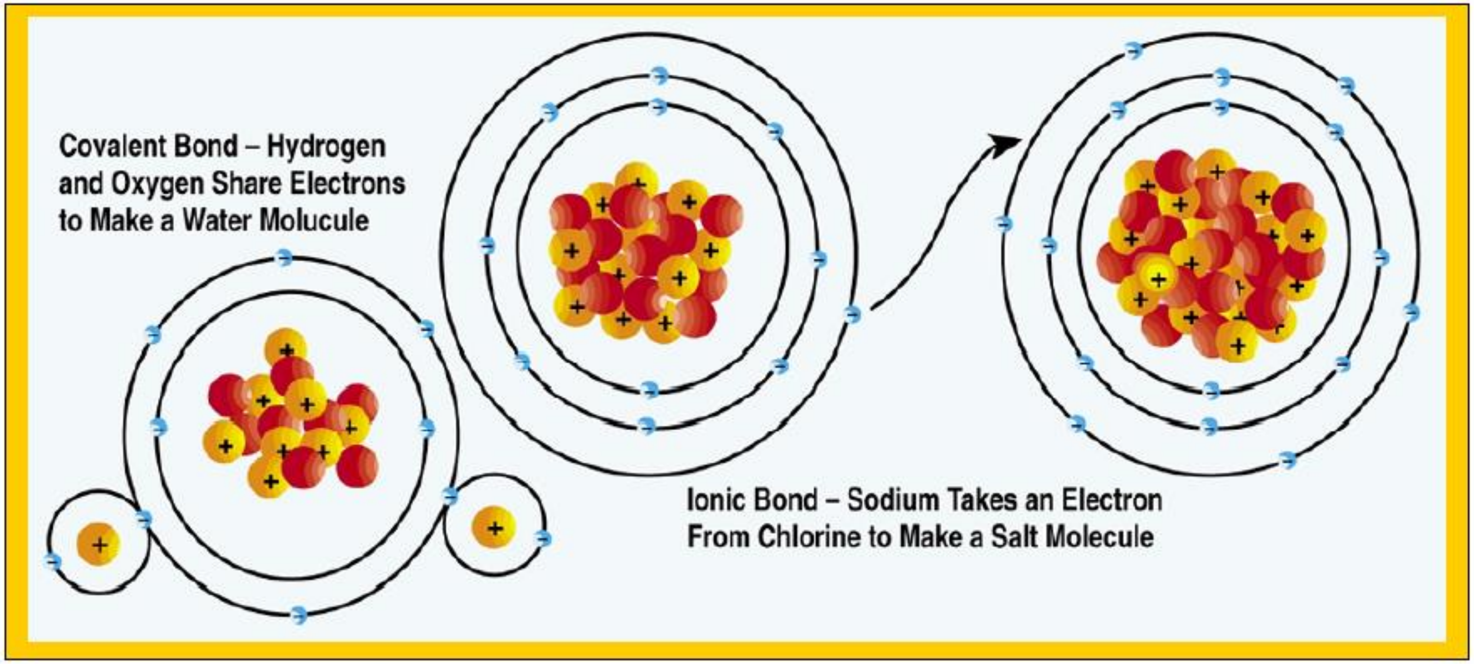
COVALENT BOND in Hydrogen (H_2)

- **Hydrogen forms one covalent bond.**
- **When two hydrogen atoms are joined in a bond, each has a filled valence shell of two electrons.**



Ionic and Covalent Bonds

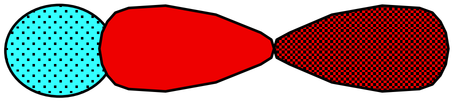
Ionic and Covalent Bonds:



Model of Covalent and Ionic Bonds

Bond Formation

A bond can result from a "head-to-head" overlap of atomic orbitals on neighboring atoms.

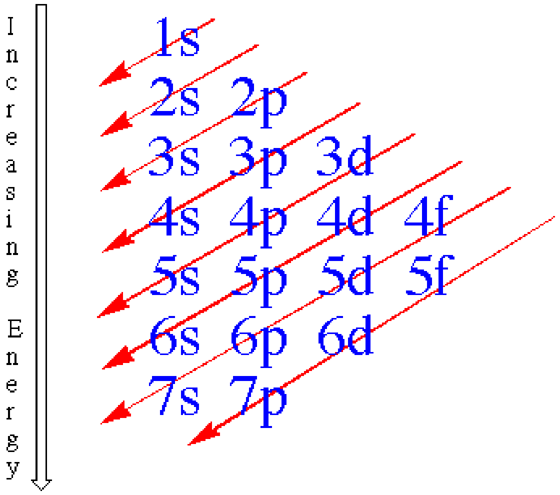


Overlap of H (1s) and Cl (2p)

This type of overlap places bonding electrons in a MOLECULAR ORBITAL along the line between the two atoms and forms a SIGMA BOND (σ).

Electronic Configuration

What is Electronic Configuration of Na (11), Mg (12) and Cl (17) atoms

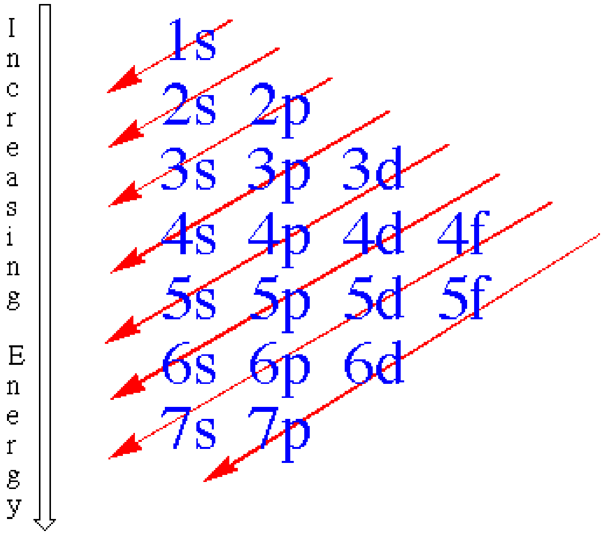


1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s

Cl = $1s^2 2s^2 2p^6 3s^2 3p^5$ 2 + 2 + 6 + 2 + 5 = 17 e

(17)

Electronic Configuration



Mg = (12) $Is^2, 2s^2, 2p^6,3s^2$ **2 + 2 + 6 + 2 = 12 electrons**

Na = (11) $Is^2, 2s^2, 2p^6,3s^1$ **2 + 2 + 6 + 1 = 11 electrons**

Structure and Bonding

- **Second row elements** can have no more than eight electrons around them. For neutral molecules, this has two consequences:
  - **Atoms with one, two, or three valence electrons form one, two, or three bonds, respectively, in neutral molecules.**
  - **Atoms with four or more valence electrons form enough bonds to give an octet. This results in the following equation:**

$$\boxed{\text{predicted number of bonds}} = 8 - \text{number of valence electrons}$$

  - **When second-row elements form fewer than four bonds their octets consist of both bonding (shared) and nonbonding (unshared) electrons. Unshared electrons are also called *lone pairs*.**

Electronic Contribution in Molecule

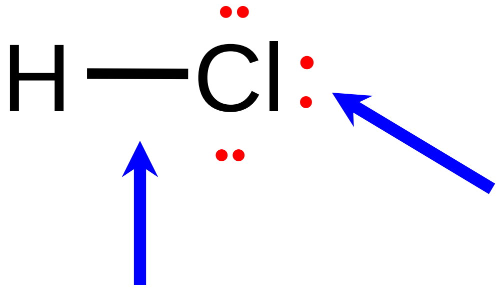
- Electron distribution is depicted with **Lewis electron dot structures**

Electrons are distributed as:

- shared or **BOND PAIRS** and
- unshared or **LONE PAIRS**.

Bond and Lone Pairs

- Electrons are distributed as shared or **BOND PAIRS** and unshared or **LONE PAIRS**.



Unshared or
lone pair (LP)

shared or bond pair

This is a LEWIS ELECTRON DOT structure.

Rule for Lewis Dot Structure:

1. Identify the **central atom** (less electronegative atom) of molecule.
   - Connect the central atom with other atoms with single bonds.

2. Calculate the number of the **electron in pi-bonds** (multiple bonds) using formula **1** as given below:

   P = 6n + 2 – V ------------------------------- (**1**)

   where as P is the number of electrons in pi-bond

   n is the number of atom in the molecule

   & V is the valence electrons in the molecule = (Sum of atom group numbers) – Charge

3. Add **double or triple bonds** to the structure in Step 1 according to the results obtained in Step 2.

   - Add unshared pairs of electrons around each atom so that all of them have octets around them (except for the H atom)

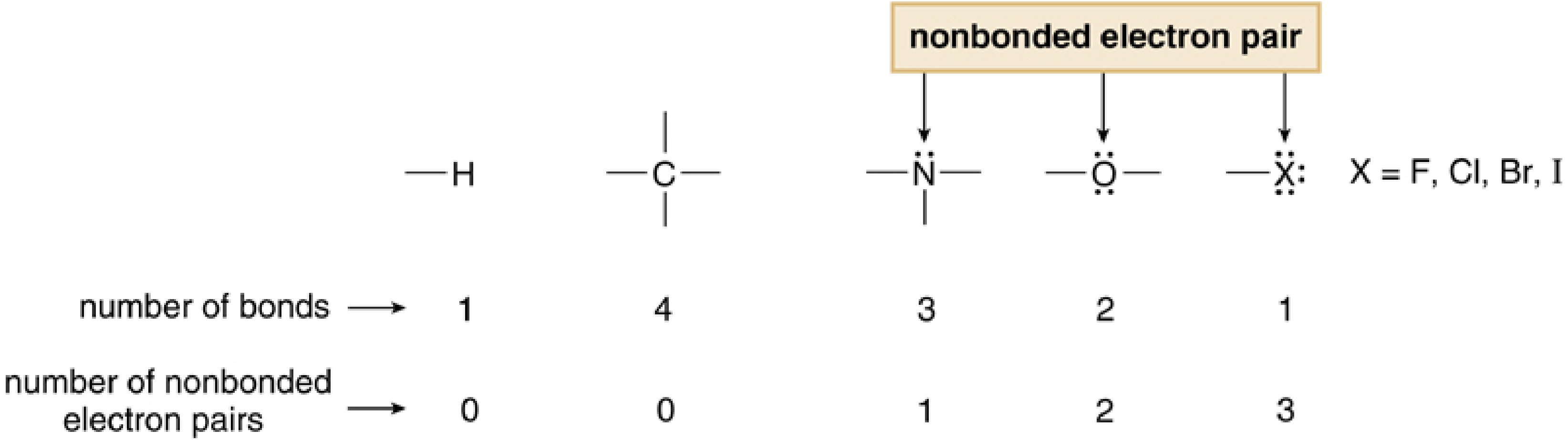
4. Calculate the **Formal Charge** of the atoms in the molecule as follows:

   Formal Charge of Element =

   Group Number – Total number of electrons "owned" by the atom concerned

Structure and Bonding

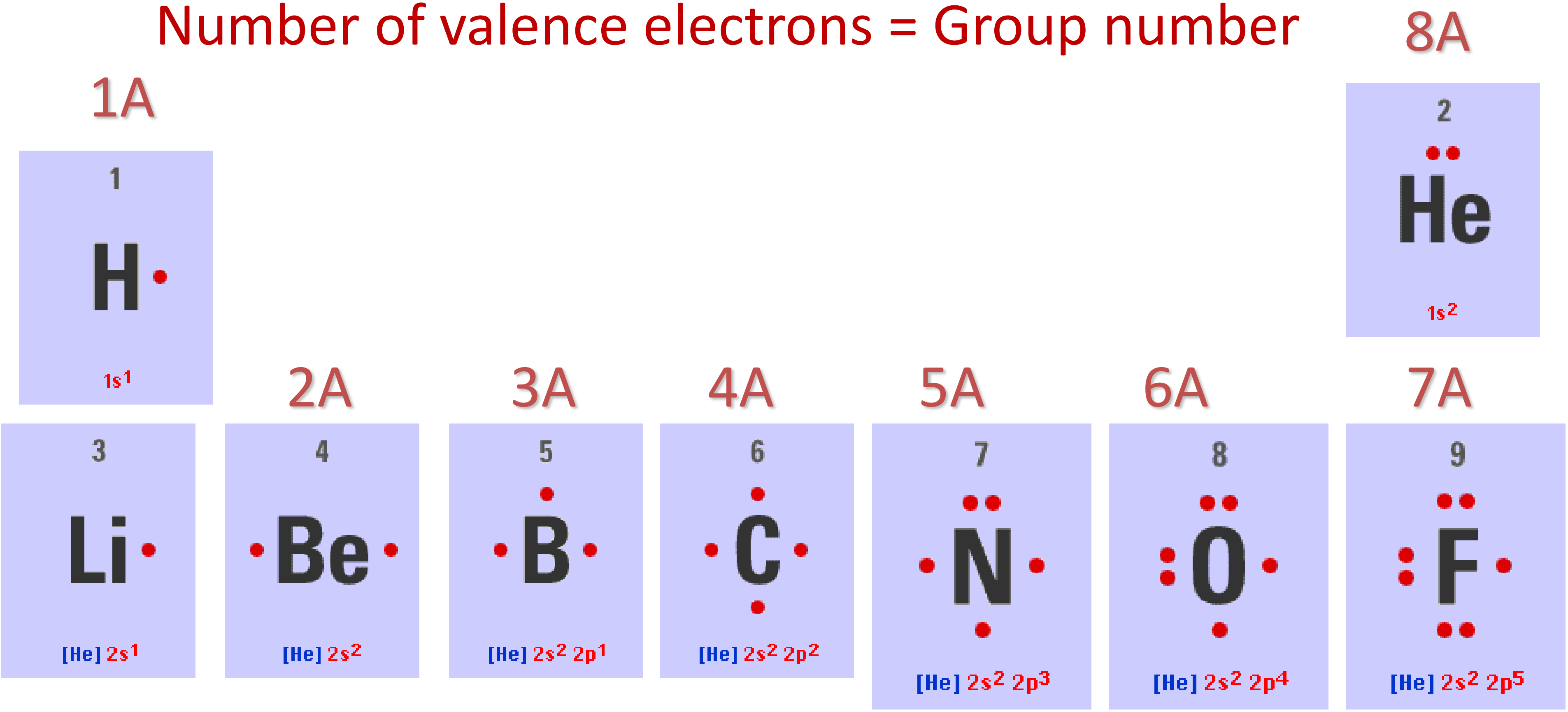
Summary: The usual number of bonds of common neutral atoms



	H	C	N	O	X
number of bonds →	1	4	3	2	1
number of nonbonded electron pairs →	0	0	1	2	3

Valence Electrons

Number of valence electrons = Group number



8A

2 He $1s^2$

1A

1 H $1s^1$

2A 3A 4A 5A 6A 7A

3 Li [He] $2s^1$

4 Be [He] $2s^2$

5 B [He] $2s^2 2p^1$

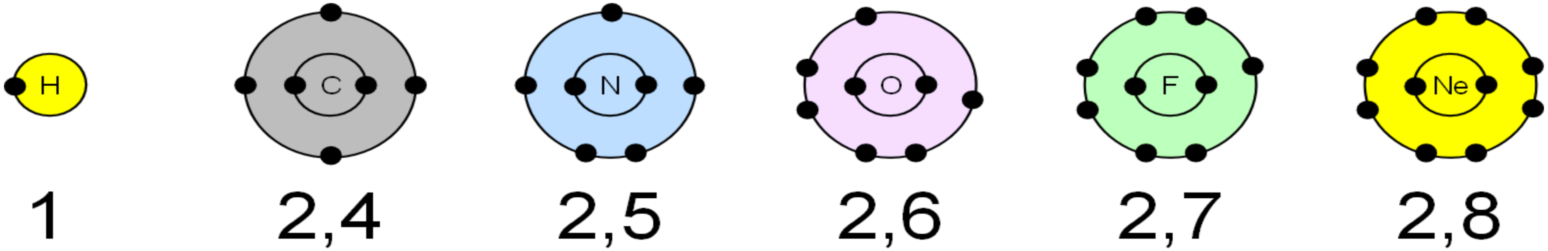
6 C [He] $2s^2 2p^2$

7 N [He] $2s^2 2p^3$

8 O [He] $2s^2 2p^4$

9 F [He] $2s^2 2p^5$

How Covalent bonds Form



The process of sharing electrons is known as covalent bonding.

In order to achieve a full outer shell...

Hydrogen atoms each need one electron;

Carbon atoms each need four electrons;

Nitrogen atoms each need three electrons;

Oxygen atoms each need two electrons;

Fluorine atoms each need one electron;

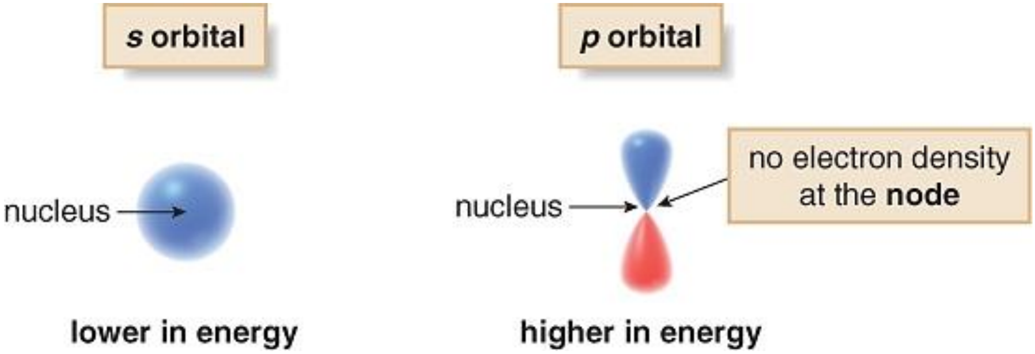
Neon atoms already have a full outer electron shell.

Structure and Bonding

TYPES OF COVALENT BOND or Molecular Orbitals

- An *sigma* orbital has a sphere of electron density and is lower in energy than the other orbitals of the same shell.
- A *p* orbital has a dumb bell shape and contains a node of electron density at the nucleus. It is higher in energy than an *s* orbital.

Sigma & pi-bond: *overlapping of orbitals.



Structure and Bonding

Since there is only one orbital in the first shell, and each shell can hold a maximum of two electrons, there are two elements in the first row, H and He.

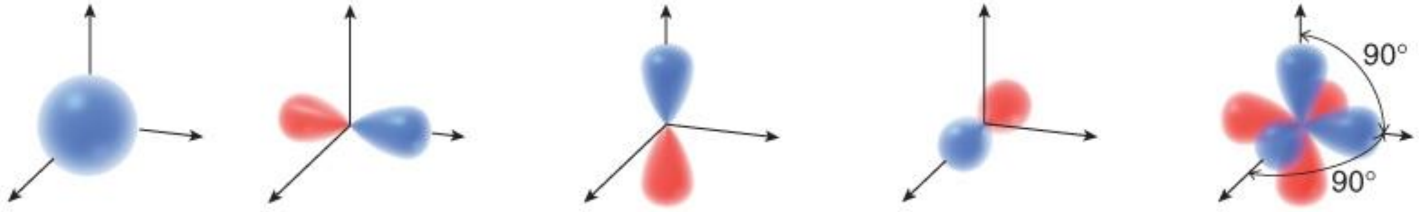
First row → H

He

$1s^1$ ← **electronic configuration** → $1s^2$

Each element in the second row of the periodic table has four orbitals available to accept additional electrons: *one 2s orbital*, and *three 2p orbitals*.

The four orbitals in the second shell



$2s$ orbital

$2p_x$ orbital

$2p_y$ orbital

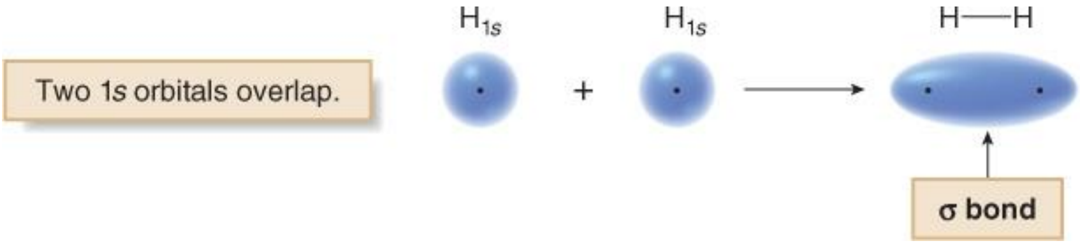
$2p_z$ orbital

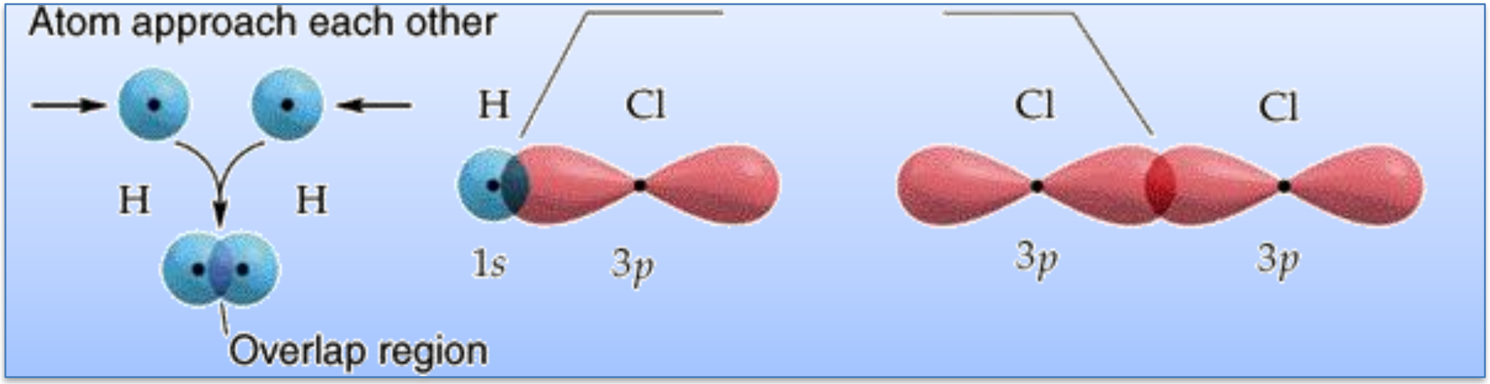
All three $2p$ orbitals drawn on the same set of axes.

Structure and Bonding

Types of sigma bonds:

1. s-s
2. s-p
3. p-p
4. sp^3



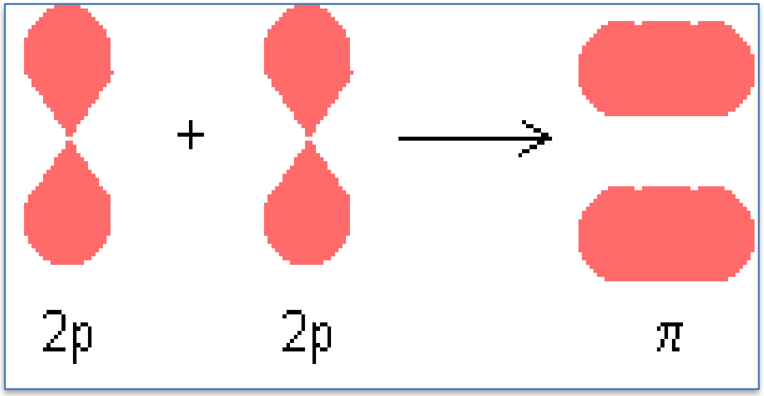


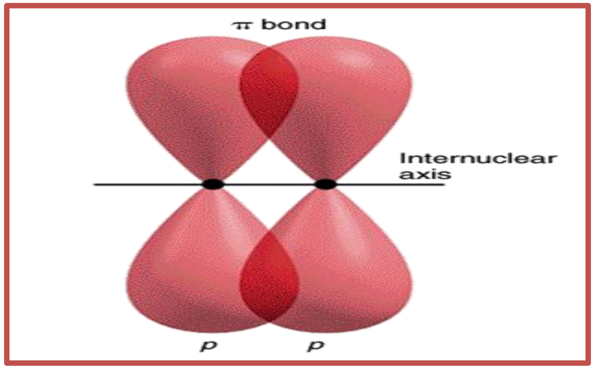
*All single bond is sigma bond.

Structure and Bonding

Pi- bond:

- The second kind of bonding is pi (π) bonding. It occurs when p orbitals overlap side by side or parallel to each other.





The increased electron density is concentrated above and below the nuclear axis.

bonding occurs in molecules with multiple bonds.

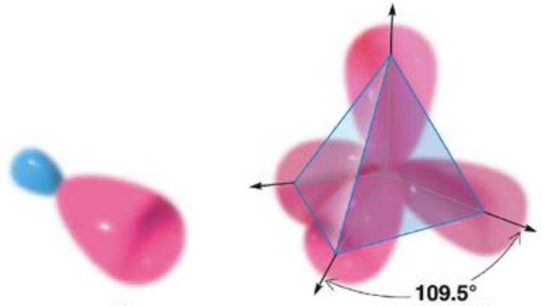
Structure and Bonding

Electronic Configration of carbon in its ground or atomic state

C (Ground state) = $1s^2\ \underbrace{2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^0}_{\text{Valence shell}}$ OR C (Ground state) = $1s^2\ 2s^2\ 2p^2$

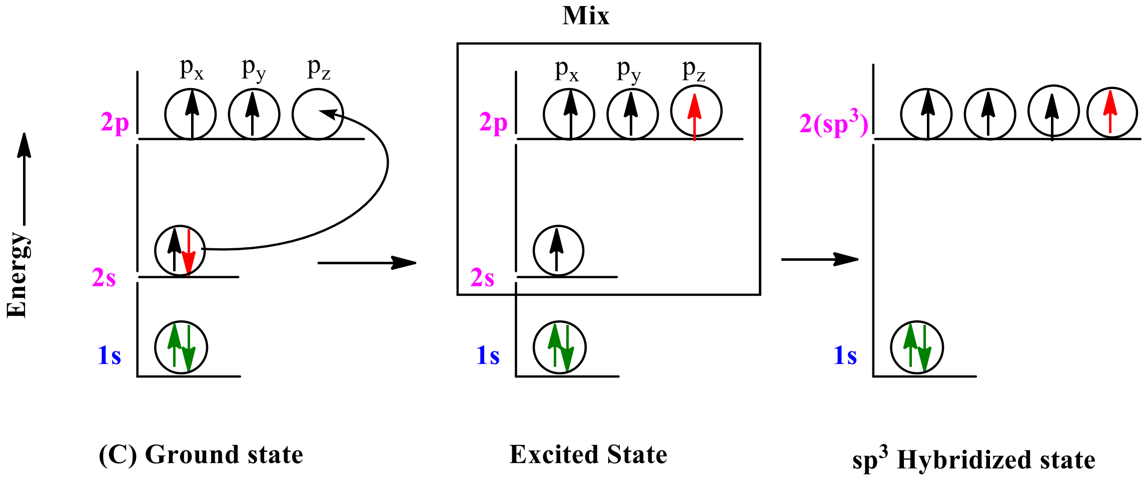
C (Excited state) = $1s^2\ \underbrace{2s^1\ 2p_x^1\ 2p_y^1\ 2p_z^1}_{\text{Valence shell}}$

C (Hybridized state) = $1s^2\ \underbrace{2(sp^3)^1\ 2(sp^3)^1 2(sp^3)^1 2(sp^3)^1}_{\text{Valence shell}}$



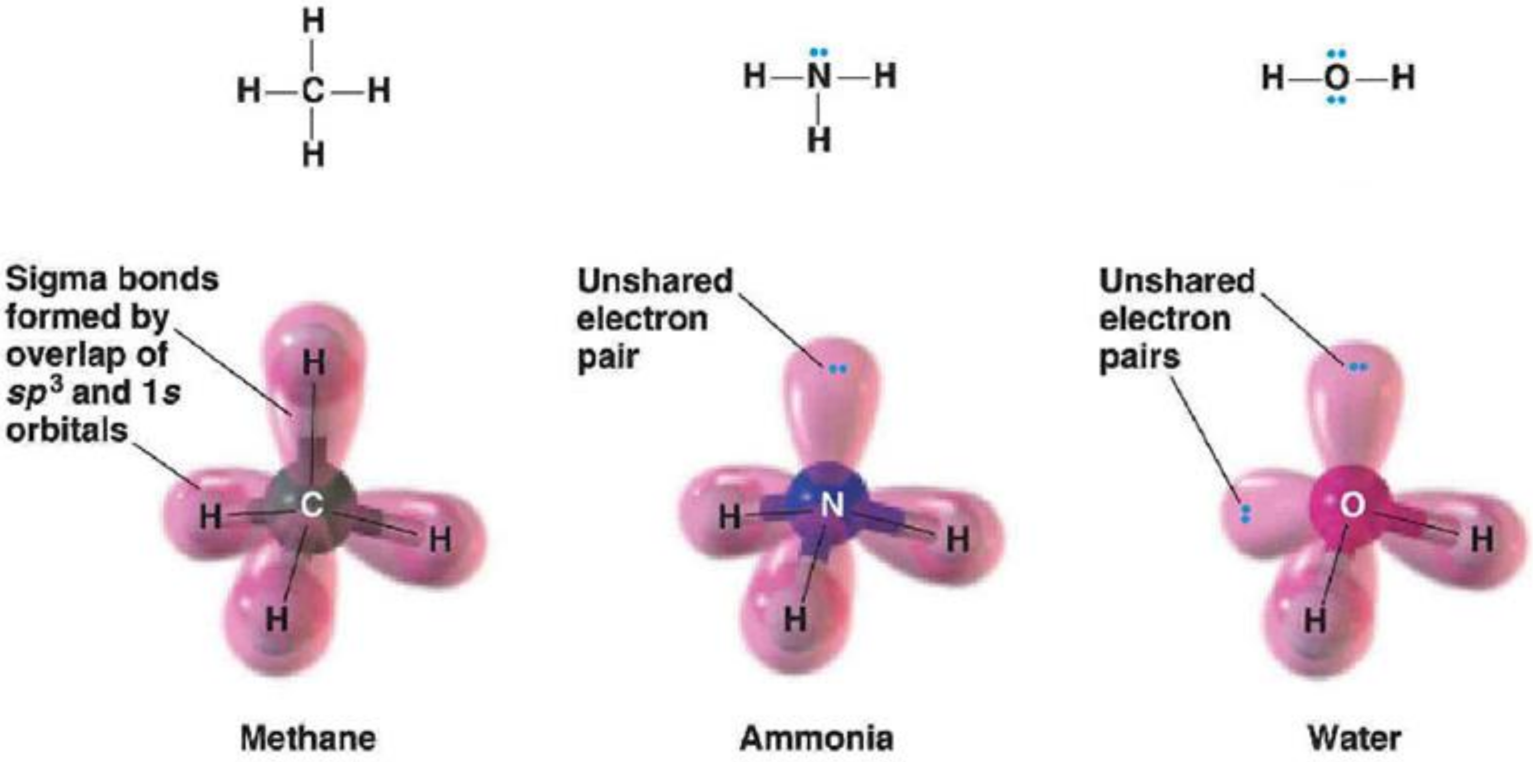
(a) An sp^3 orbital (b) Four tetrahedral sp^3 orbitals

Energy-level diagram



Structure and Bonding

orbital overlap pictures of methane, ammonia, and water



Structure and Bonding

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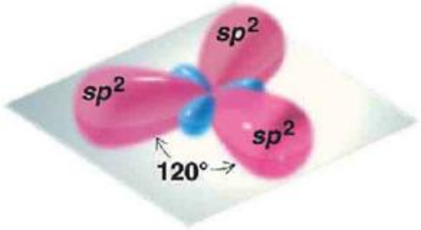
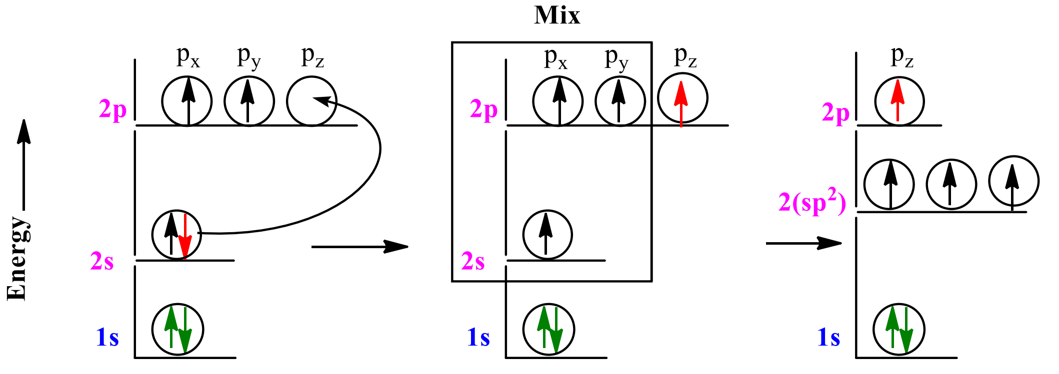
C (Hybridized state) = $1s^2\ \underbrace{2(sp^2)^1\ 2(sp^2)^1 2(sp^2)^1 2p_z^1}_{\text{Valence shell}}$

(a) An *sp*² orbital



(b) Three *sp*² orbitals

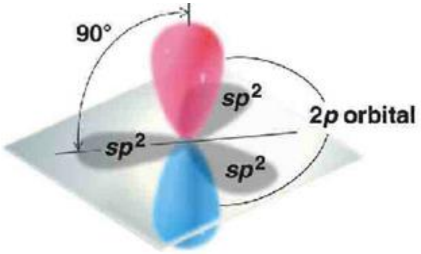
Energy-level diagram



(C) Ground state

Excited State

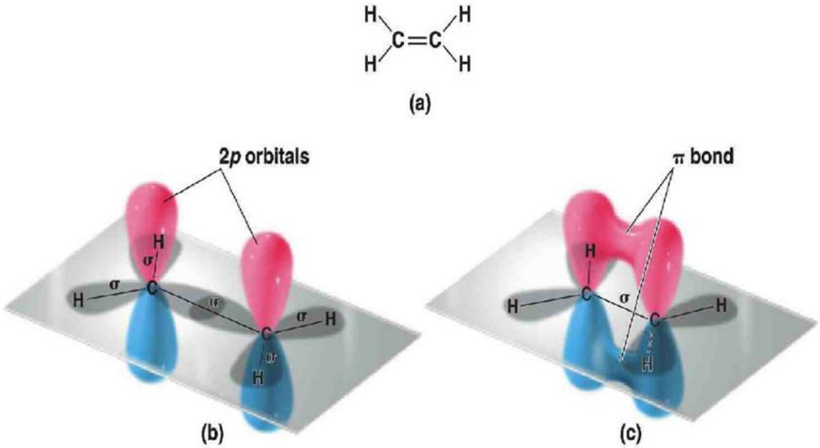
sp² Hybridized state



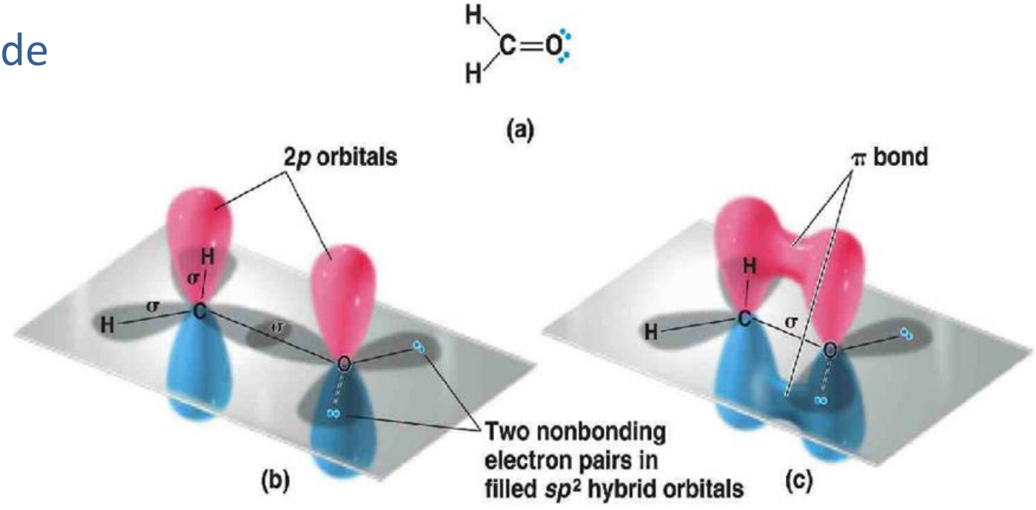
(c) Three *sp*² orbitals and an unhybridized 2*p* orbital

Structure and Bonding

Bonding in Ethylene



Bonding in Formaldehyde



Two nonbonding electron pairs in filled sp^2 hybrid orbitals

Structure and Bonding

Electronic Configration of carbon in its ground or atomic state

C (Ground state) = $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^0$ OR C (Ground state) = $1s^2\ 2s^2\ 2p^2$

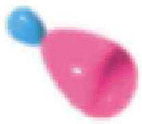
Valence shell

C (Excited state) = $1s^2\ 2s^1\ 2p_x^1\ 2p_y^1\ 2p_z^1$

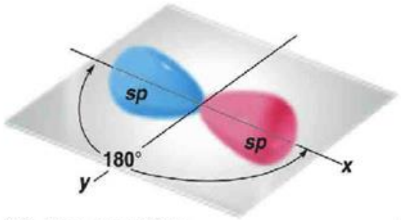
Valence shell

C (Hybridized state) = $1s^2\ 2(sp)^1\ 2(sp)^1 2p_y^1 2p_z^1$

Valence shell

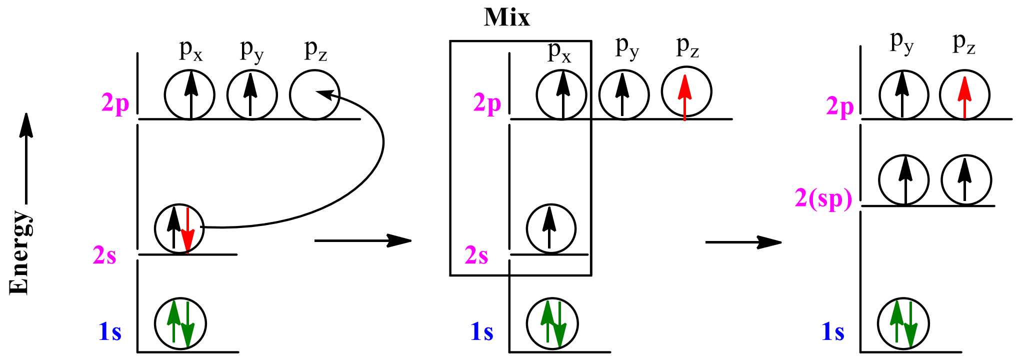


(a) An *sp* orbital

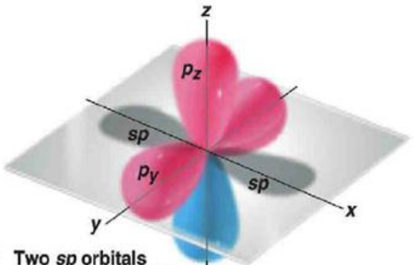


(b) Two *sp* orbitals

Energy-level diagram formate



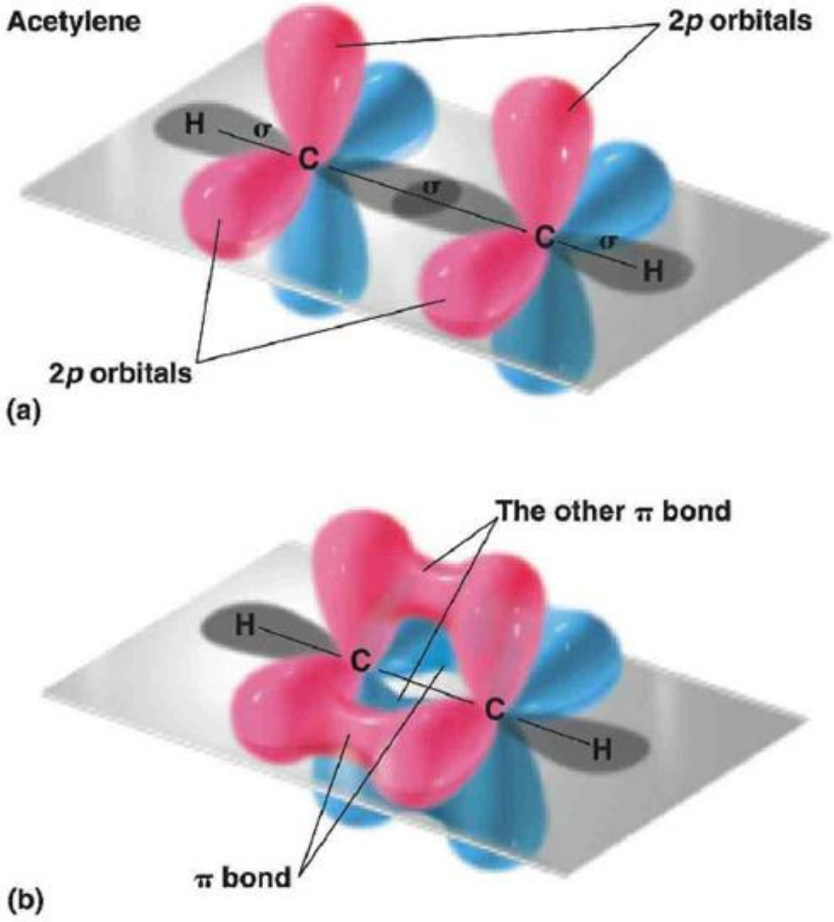
(C) Ground state **Excited State** **sp Hybridized state**



(c) Two *sp* orbitals and two unhybridized 2*p* orbitals

Structure and Bonding

Bonding in Acetylene



Structure and Bonding

σ-bond
Single bond = σ bond
Double bond = $\sigma + \pi$

π-bond
(weaker)
Triple bond = $\sigma + 2\pi$

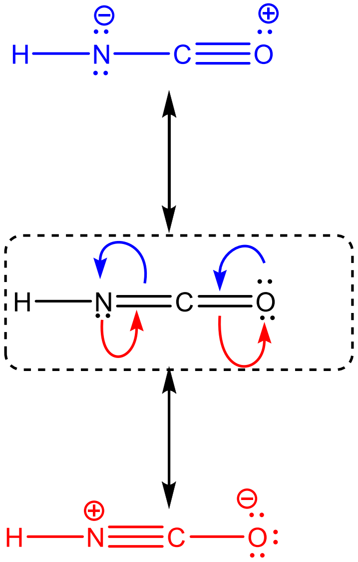
e.g.

H-H	CO_2	N_2
F-F	CS_2	HCN
H-F		
H_3C-CH_3		

Resonances

Resonances: occurs whenever a molecule can be represented by two or more structures differing only in the arrangement of electrons, without shifting any atoms. It only involves the delocalization of electrons.

When one or more valid Lewis dot diagram can be drawn. e.g. HNCO and SO_3



no change in

1. net # of e
2. net charge
3. sigma bond (connectivity)

Resonance contribution

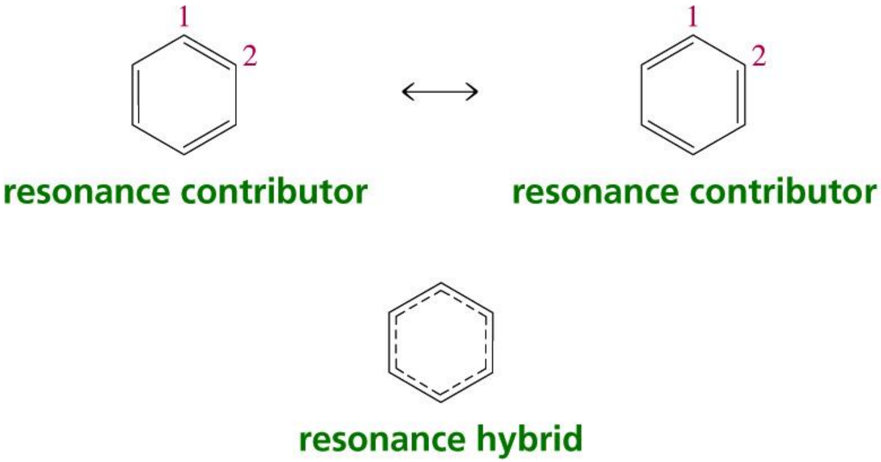
1. closed shells
2. minimize Formal charge
3. If F.C. is present then
   (-) with electronegative atoms
   (+) w/o electronegative atoms

Resonance Hybrid

- A compound with delocalized e$^-$ is said to have ***resonance***
  - resonance contributor
  - resonance structure
  - contributing resonance structure

Resonance Hybrid

- Benzene
  - – contributing resonance structures



Resonance Hybrids

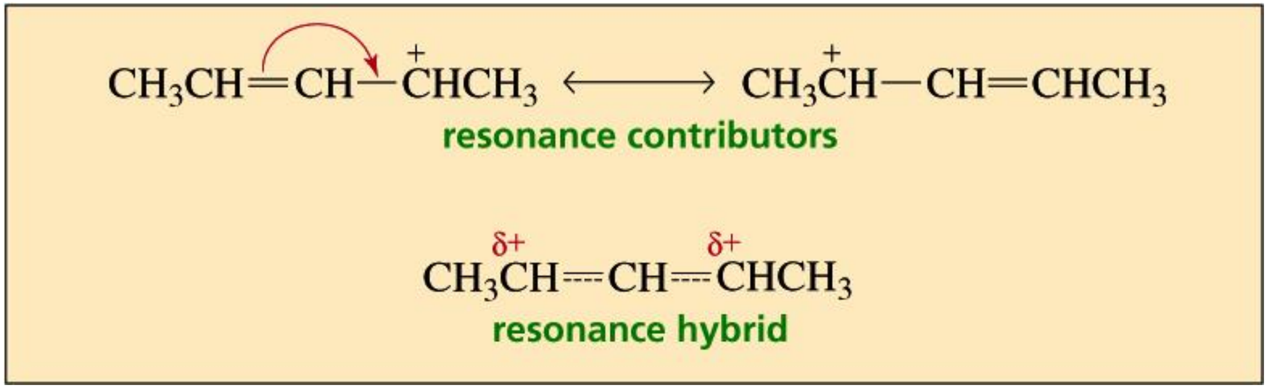
Drawing resonance hybrids

- **Only e^- move** (not atoms)
- **Only π and non-bonding e^- move**
- **Total # e^- stays same** (as does unpaired e^-)

Resonance Hybrids

e^- can be moved only by…

π e^- move toward + or toward π bond



$CH_3CH{=}CH-\overset{+}{C}HCH_3 \longleftrightarrow CH_3\overset{+}{C}H-CH{=}CHCH_3$

resonance contributors

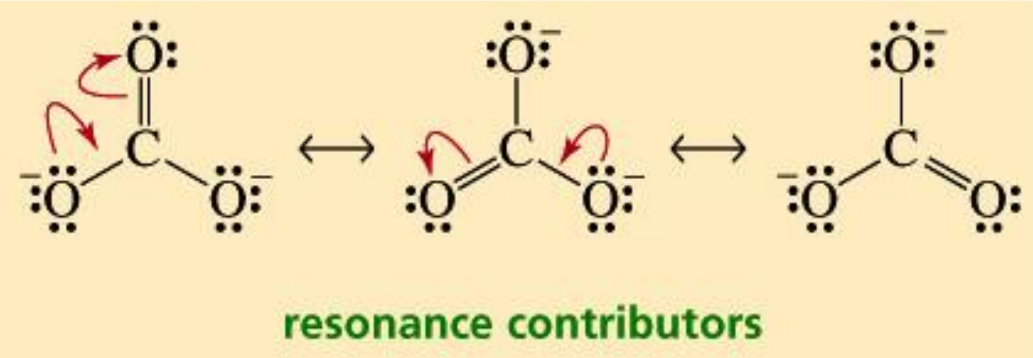
$CH_3\overset{\delta+}{C}H{=}CH{=}\overset{\delta+}{C}HCH_3$

resonance hybrid

Resonance Hybrids

*e^- can be moved **only** by…*

Nonbonding pair e^- toward a π bond

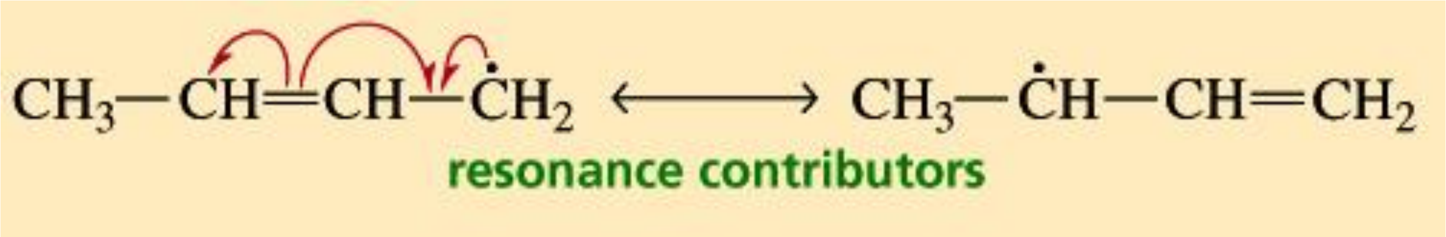


resonance contributors

Resonance Hybrids

*e⁻ can be moved **only** by…*

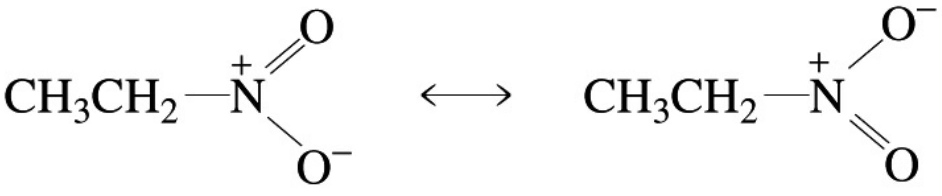
Nonbonding single e⁻ toward a π bond



$CH_3{-}CH{=}CH{-}\dot{C}H_2 \longleftrightarrow CH_3{-}\dot{C}H{-}CH{=}CH_2$

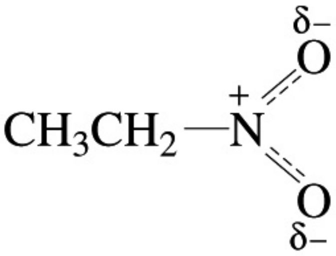
Resonance Hybrids

- Drawing resonance hybrids of nitro ethane



resonance contributor

resonance contributor



resonance hybrid

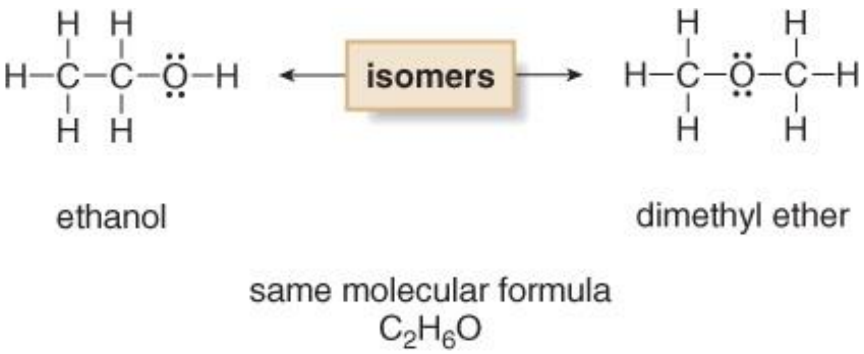
Isomers

Isomers

Isomers have same molecular formula but different molecular structure called isomer.

In drawing a Lewis structure for a molecule with several atoms, sometimes more than one arrangement of atoms is possible for a given molecular formula.

Example:



Both are valid Lewis structures and both molecules exist. These two compounds are called isomers.

Ethanol and dimethyl ether are constitutional isomers.

Isomers

Examples of isomers:

$CH_2{=}\ CH\text{-}CH_2\text{-}CH_3$

I-butene

C_4H_8

$CH_3\text{-}CH{=}CH\text{-}CH_3$

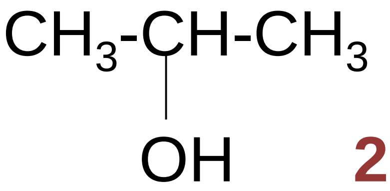
2-butene

C_4H_8

$CH_3\text{-}CH_2\text{-}CH_2\text{-}OH$

1-Propanol

C_3H_8O



$$CH_3\text{-}\underset{\displaystyle OH}{\underset{|}{C}}H\text{-}CH_3$$

2-Propanol

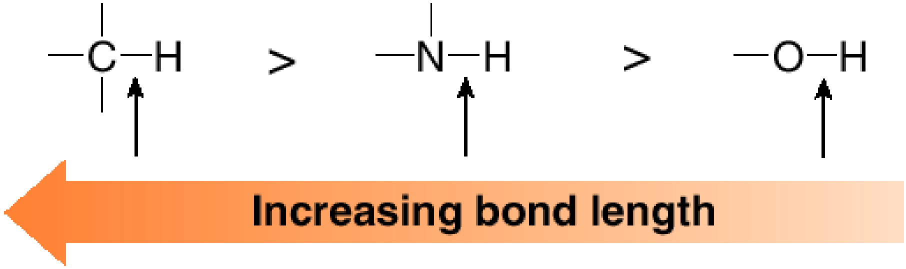
C_3H_8O

Structure and Bonding

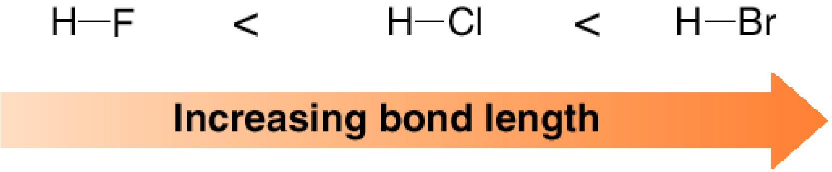
Determining Molecular Shape

Two variables define a molecule’s structure: *bond length* and *bond angle*.

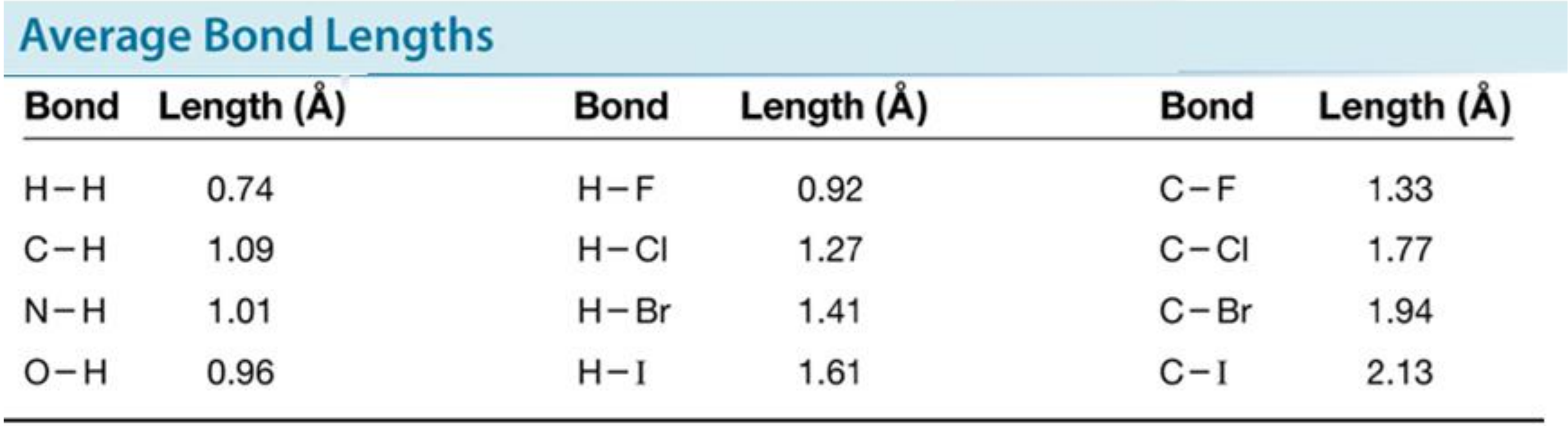
- Bond length *decreases* across a row of the periodic table as the size of the atom *decreases*.



- Bond length *increases* down a column of the periodic table as the size of an atom *increases*.



Structure and Bonding



Average Bond Lengths

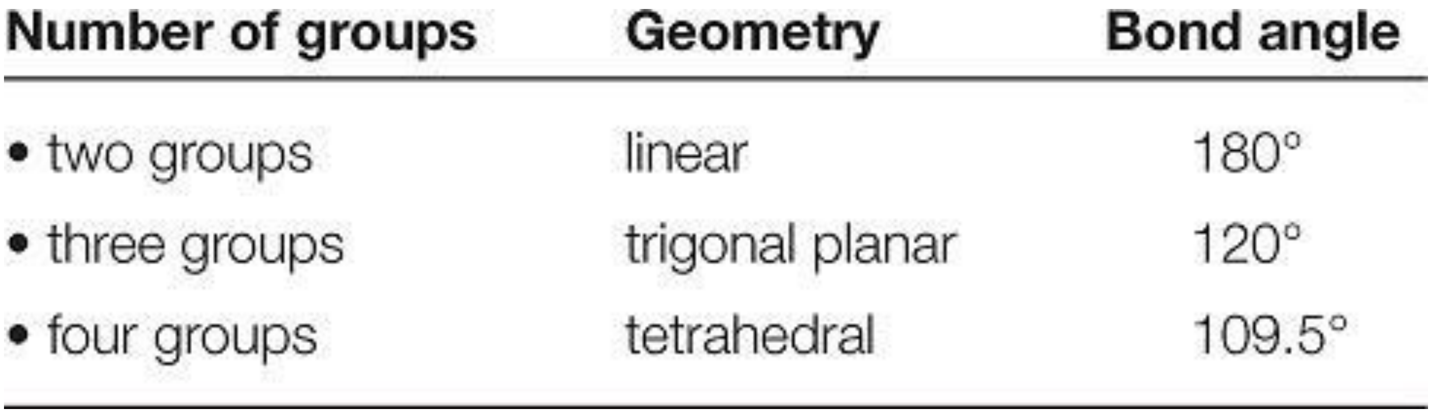
Bond	Length (Å)	Bond	Length (Å)	Bond	Length (Å)
H−H	0.74	H−F	0.92	C−F	1.33
C−H	1.09	H−Cl	1.27	C−Cl	1.77
N−H	1.01	H−Br	1.41	C−Br	1.94
O−H	0.96	H−I	1.61	C−I	2.13

Structure and Bonding

Determining Molecular Shape—Bond Angle

Bond angle determines the shape around any atom bonded to two other atoms.

- The number of groups surrounding a particular atom determines its geometry. A group is either an atom or a lone pair of electrons.

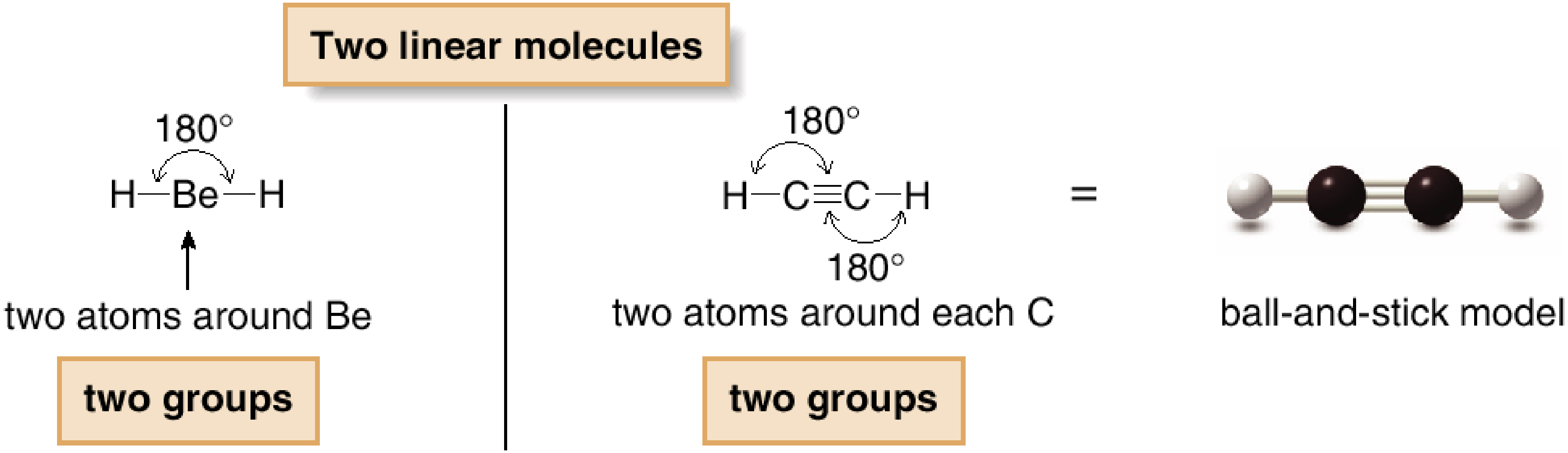


Number of groups	**Geometry**	**Bond angle**
• two groups	linear	180°
• three groups	trigonal planar	120°
• four groups	tetrahedral	109.5°

Structure and Bonding

Determining Molecular Shape—Bond Angle

Two groups around an atom—



Two linear molecules

180°

$H-Be-H$

two atoms around Be

two groups

180°

$H-C\equiv C-H$

180°

two atoms around each C

two groups

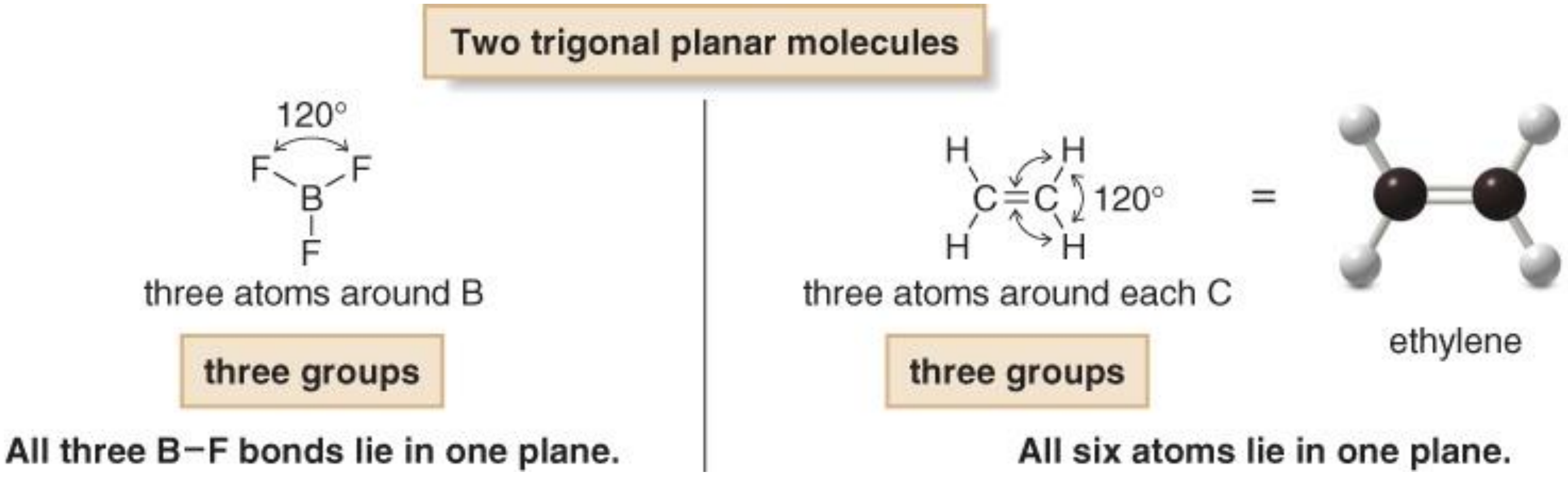
=

ball-and-stick model

Structure and Bonding

Determining Molecular Shape—Bond Angle

Three groups around an atom—



All three B–F bonds lie in one plane.

All six atoms lie in one plane.

Structure and Bonding

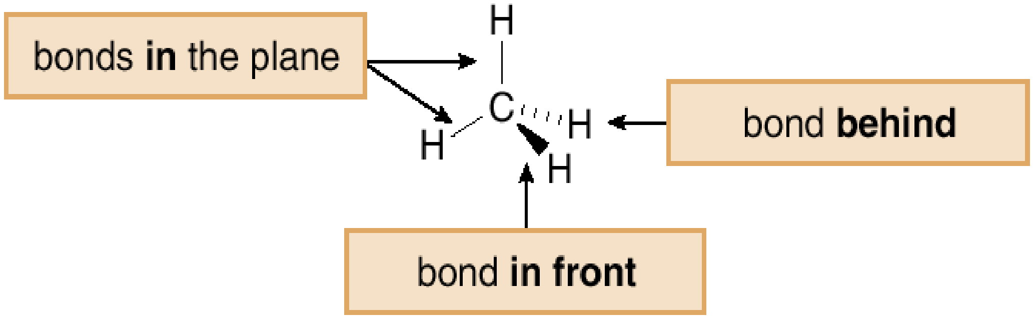
Determining Molecular Shape

Four groups around an atom—

Structure and Bonding

Drawing Three Dimensional Structures

- A solid line is used for a bond in the plane.
- A wedge is used for a bond in front of the plane.
- A dashed line is used for a bond behind the plane.



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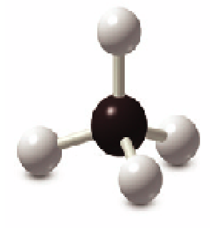
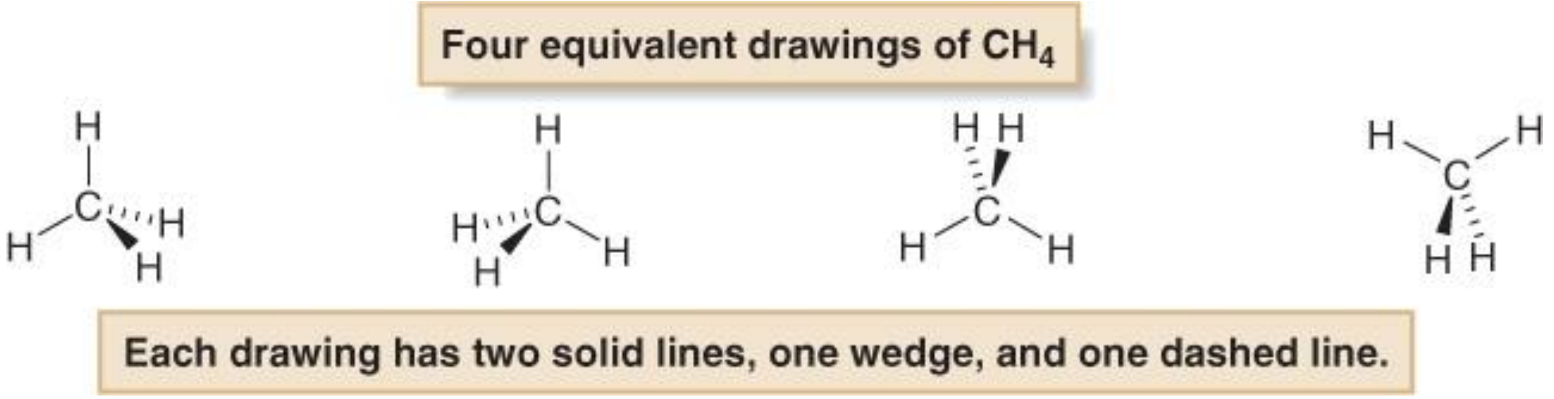


ball-and-stick model of CH_4

Structure and Bonding

Drawing Three Dimensional Structures

The molecule can be turned in many different ways, generating many equivalent representations. All of the following are acceptable drawings for CH_4.



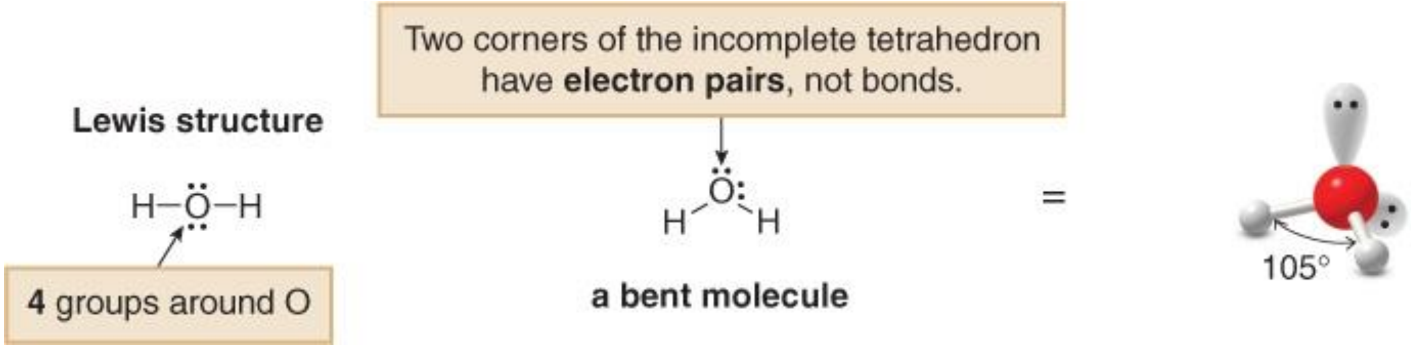
Four equivalent drawings of CH_4

Each drawing has two solid lines, one wedge, and one dashed line.

Structure and Bonding

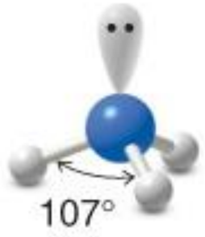
A Nonbonded Pair of Electrons is Counted as a "Group"

In water (H_2O), two of the four groups attached to the central O atom are lone pairs. The two H atoms and two lone pairs around O point to the corners of a tetrahedron. The H-O-H bond angle of 105° is close to the theoretical tetrahedral bond angle of 109.5°. Water has a bent shape, because the two groups around oxygen are lone pairs of electrons.

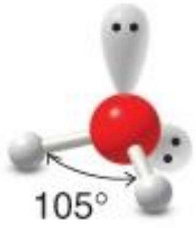


Structure and Bonding

Methane (CH_4)



Ammonia (NH_3)

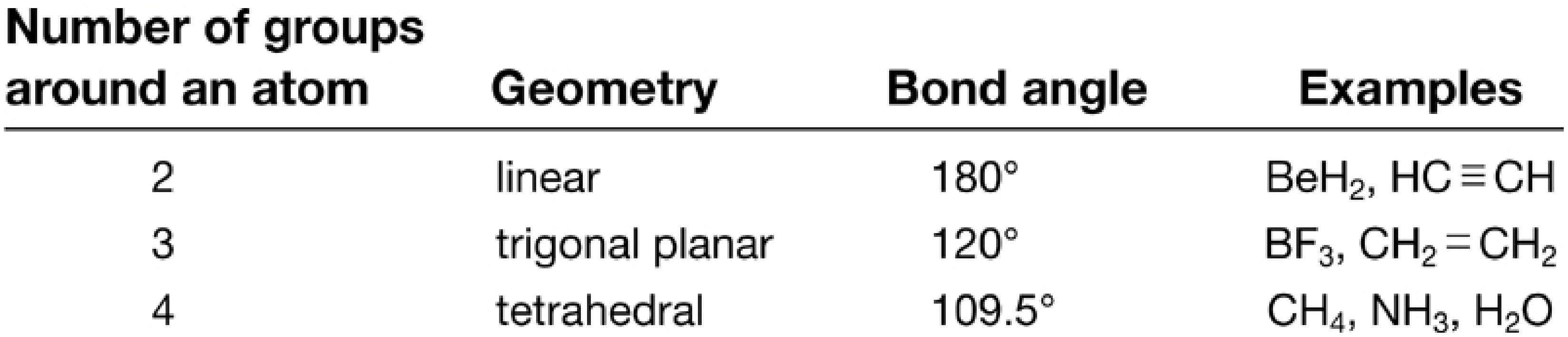


Water (H_2O)

In both NH_3 and H_2O, the bond angle is smaller than the theoretical tetrahedral bond angle (109.5°) because of repulsion of the lone pairs of electrons. The bonded atoms are compressed into a smaller space with a smaller bond angle.

Structure and Bonding

Predicting Geometry Based on Counting of Groups Around the Central Atom

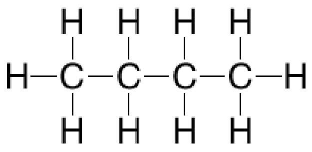


Number of groups around an atom	Geometry	Bond angle	Examples
2	linear	180°	BeH_2, $HC\equiv CH$
3	trigonal planar	120°	BF_3, $CH_2{=}CH_2$
4	tetrahedral	109.5°	CH_4, NH_3, H_2O

Structure and Bonding

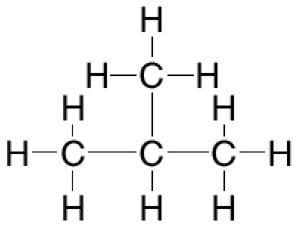
Drawing Organic Molecules—Condensed Structures

- All atoms are drawn in, but the two-electron bond lines are generally omitted.
- Atoms are usually drawn next to the atoms to which they are bonded.
- Parentheses are used around similar groups bonded to the same atom.
- Lone pairs are omitted.



= $CH_3CH_2CH_2CH_3$ or $CH_3(CH_2)_2CH_3$

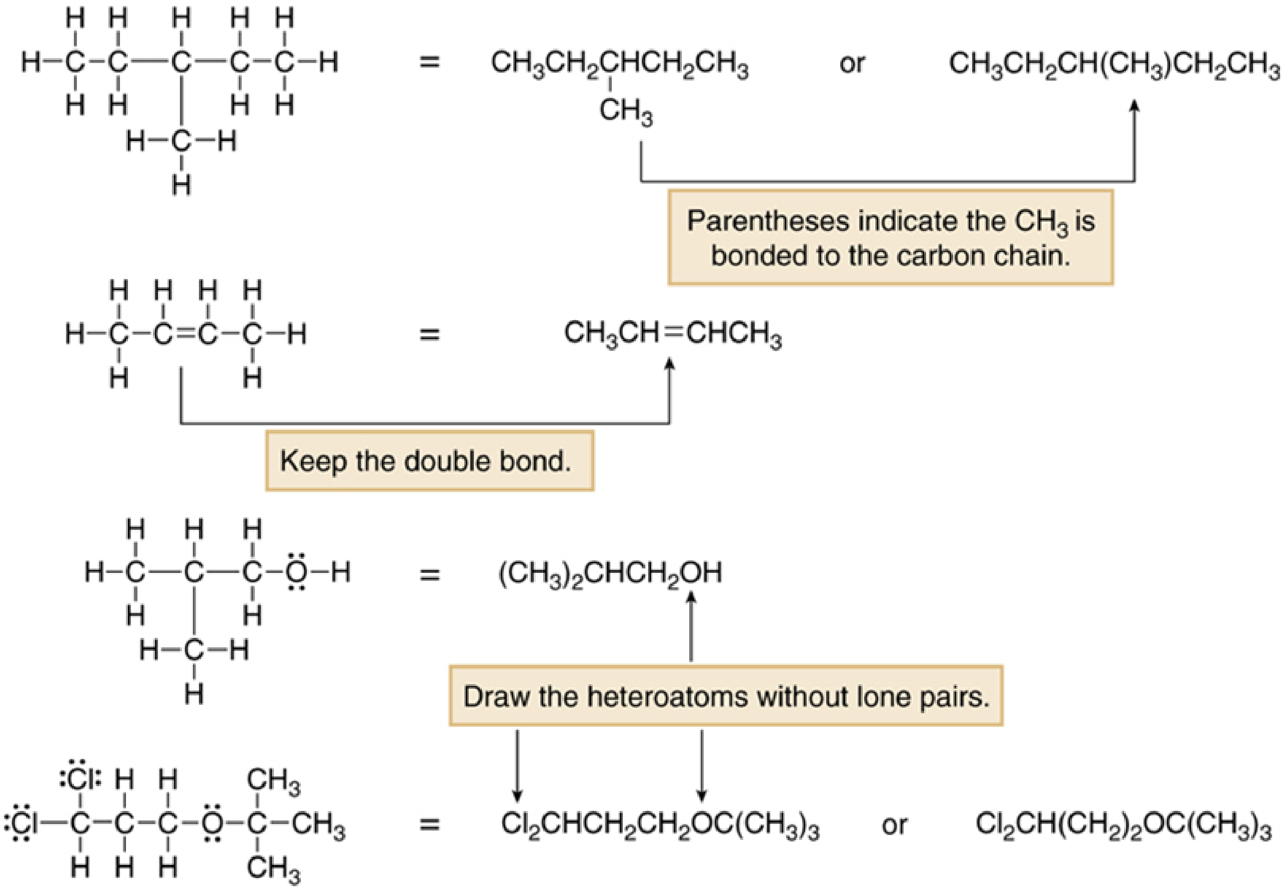
2 CH_2 groups bonded together



= $(CH_3)_3CH$

Structure and Bonding

Examples of Condensed Structures



= $CH_3CH_2\underset{\overset{|}{CH_3}}{C}HCH_2CH_3$ or $CH_3CH_2CH(CH_3)CH_2CH_3$

Parentheses indicate the CH_3 is bonded to the carbon chain.

= $CH_3CH{=}CHCH_3$

Keep the double bond.

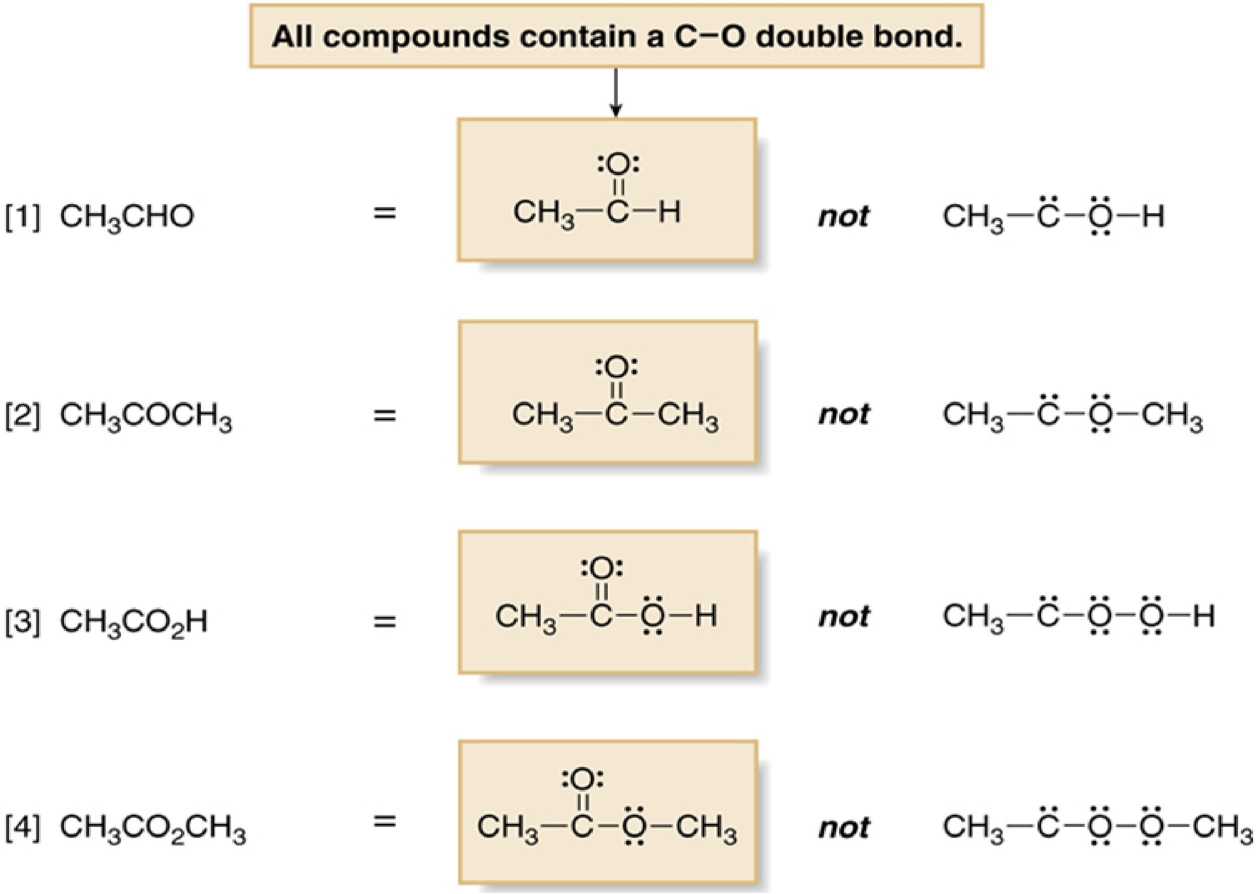
= $(CH_3)_2CHCH_2OH$

Draw the heteroatoms without lone pairs.

= $Cl_2CHCH_2CH_2OC(CH_3)_3$ or $Cl_2CH(CH_2)_2OC(CH_3)_3$

Structure and Bonding

Examples of Condensed Structures Containing a C-O Double Bond



All compounds contain a C–O double bond.

[1] CH_3CHO = $CH_3-\overset{\overset{:O:}{\|}}{C}-H$ ***not*** $CH_3-\ddot{C}-\underset{\cdot\cdot}{\ddot{O}}-H$

[2] CH_3COCH_3 = $CH_3-\overset{\overset{:O:}{\|}}{C}-CH_3$ ***not*** $CH_3-\ddot{C}-\underset{\cdot\cdot}{\ddot{O}}-CH_3$

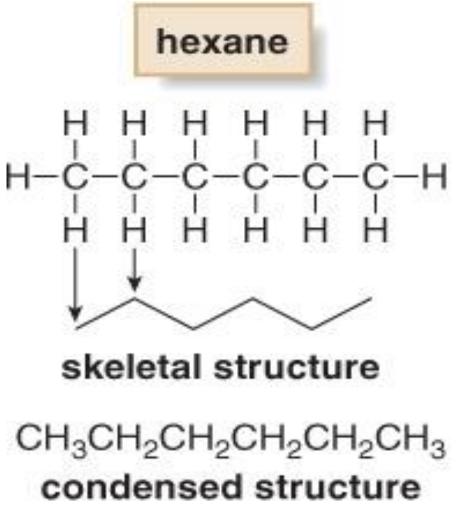
[3] CH_3CO_2H = $CH_3-\overset{\overset{:O:}{\|}}{C}-\underset{\cdot\cdot}{\ddot{O}}-H$ ***not*** $CH_3-\ddot{C}-\underset{\cdot\cdot}{\ddot{O}}-\underset{\cdot\cdot}{\ddot{O}}-H$

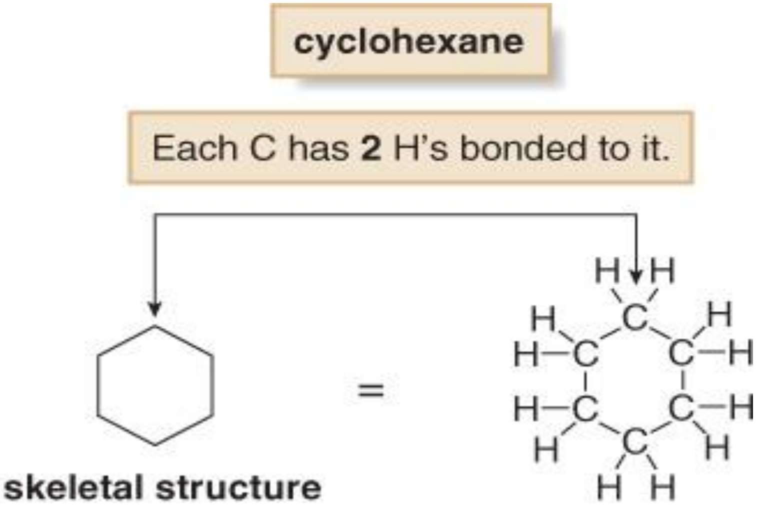
[4] $CH_3CO_2CH_3$ = $CH_3-\overset{\overset{:O:}{\|}}{C}-\underset{\cdot\cdot}{\ddot{O}}-CH_3$ ***not*** $CH_3-\ddot{C}-\underset{\cdot\cdot}{\ddot{O}}-\underset{\cdot\cdot}{\ddot{O}}-CH_3$

Structure and Bonding

Skeletal Structures

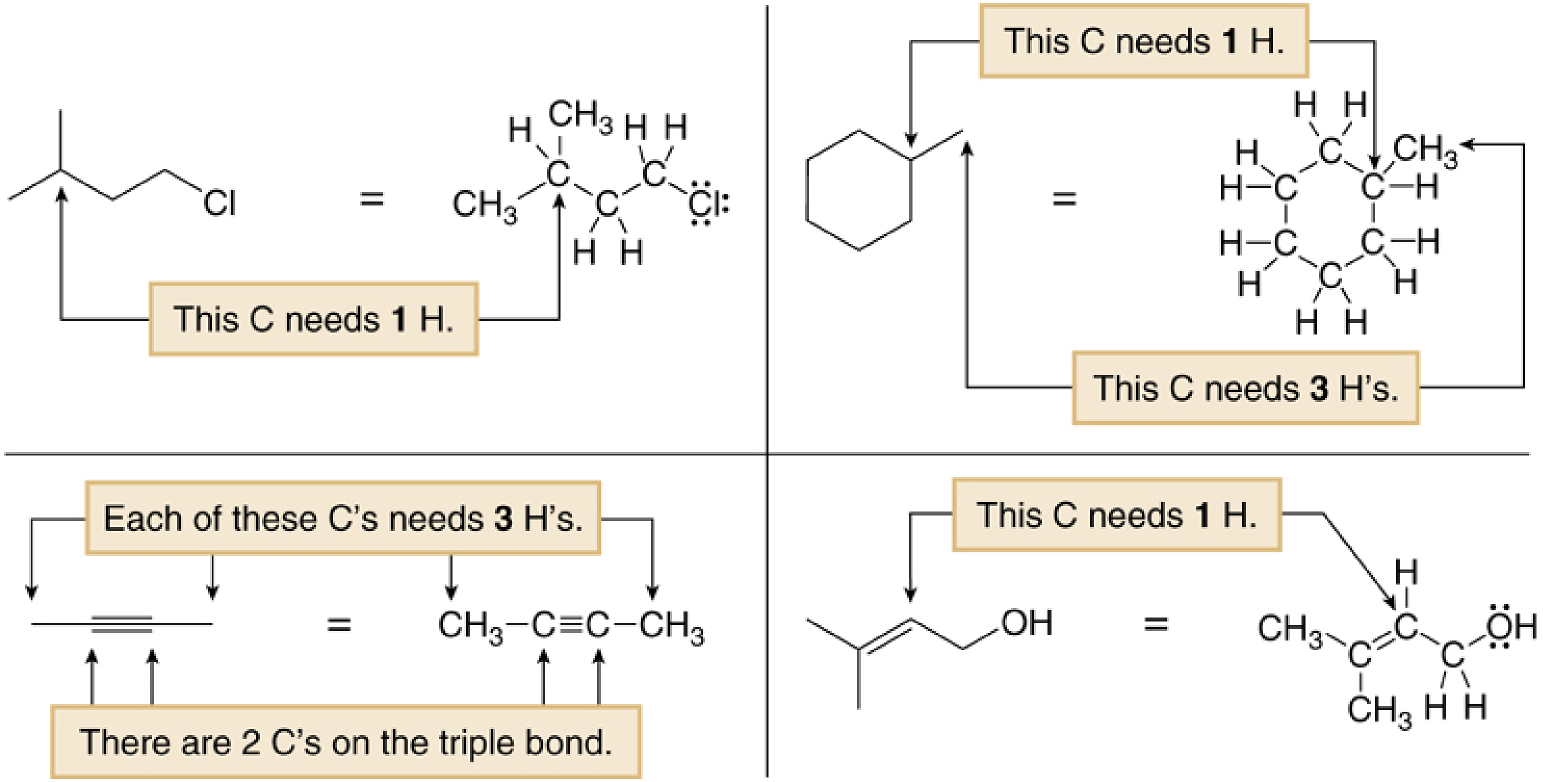
- Assume there is a carbon atom at the junction of any two lines or at the end of any line.
- Assume there are enough hydrogens around each carbon to make it tetravalent.
- Draw in all heteroatoms and hydrogens directly bonded to them.





Structure and Bonding

Examples of Skeletal Structures



This C needs **1** H.

This C needs **1** H.

This C needs **3** H's.

Each of these C's needs **3** H's.

There are 2 C's on the triple bond.

This C needs **1** H.