Kolej Matrikulasi Johor
KMJ
Keilmuan Menjana Jatidiri

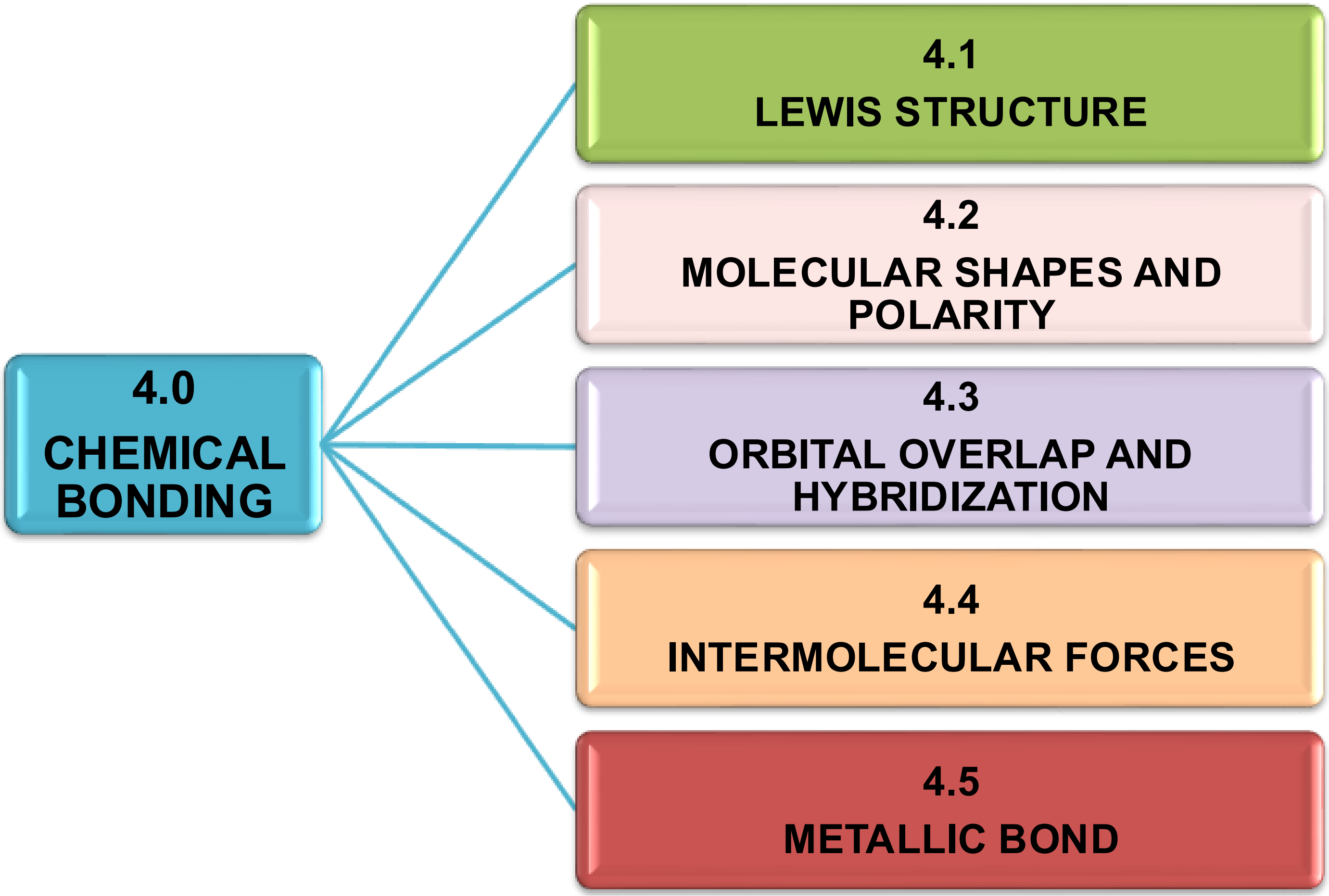
CHAPTER 4.0
CHEMICAL BONDING

CHEMISTRY 1
SK015

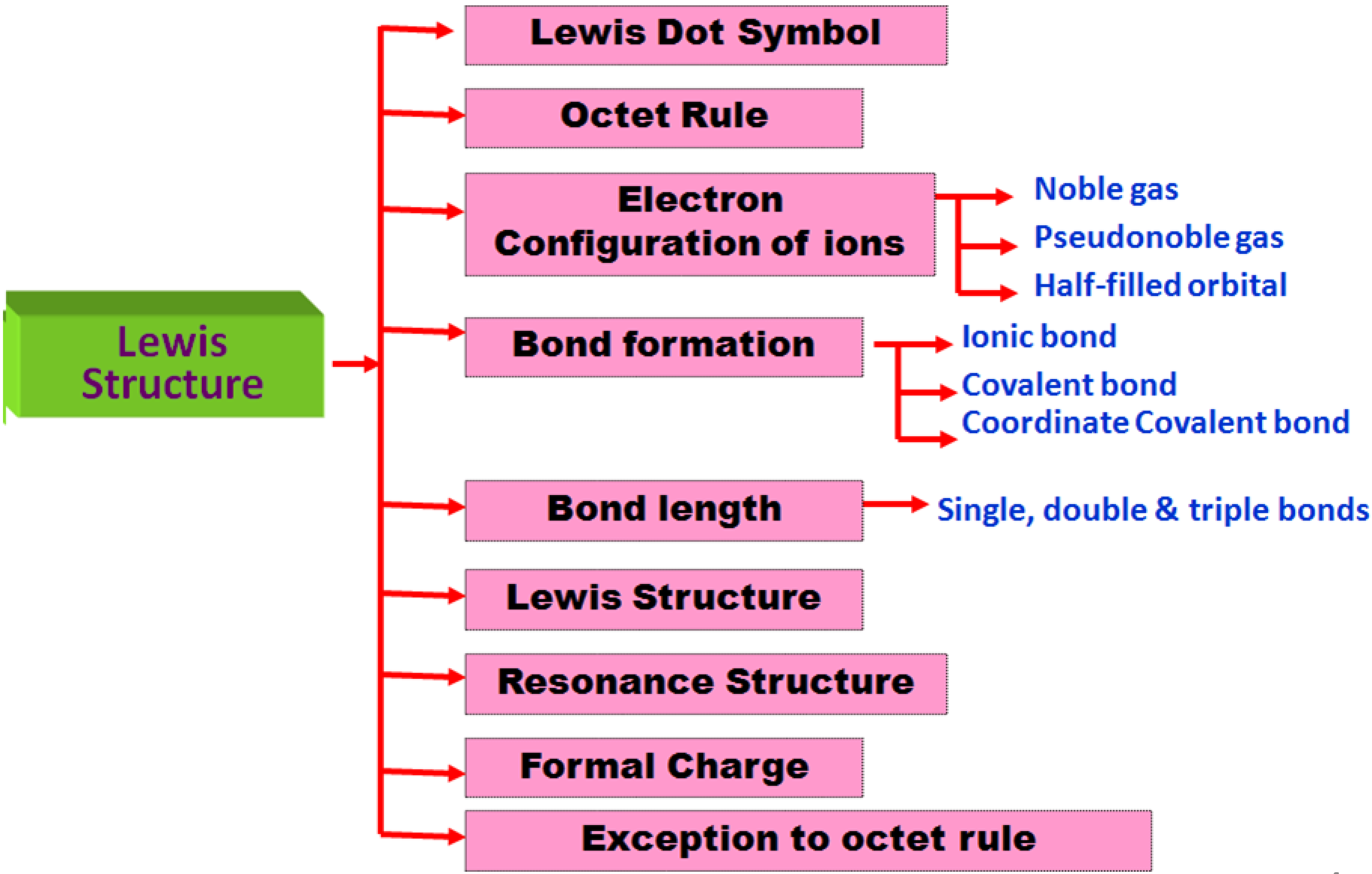
SESSION 2025/2026

STUDENT LEARNING TIME (SLT): LECTURE	
NON FACE-TO-FACE (PREPARATION)	FACE-TO-FACE (DURING CLASS)
3 HOURS	3 HOURS

CHEMISTRY UNIT, KMJ


4.0
CHEMICAL BONDING
4.1
LEWIS STRUCTURE
4.2
MOLECULAR SHAPES AND POLARITY
4.3
ORBITAL OVERLAP AND HYBRIDIZATION
4.4
INTERMOLECULAR FORCES
4.5
METALLIC BOND

CHAPTER 4.1 : OVERVIEW



4.1 LEWIS STRUCTURE

Teaching and learning outcomes

At the end of the lesson, student should be able to

4.1	**Lewis Structure**
a)	State the octet rule. (C1)
b)	Describe how atoms achieve stability by attaining stable configuration of: (C1, C2) i. Noble gas ii. Pseudo-noble gas iii. Half-filled orbital.
c)	Describe the formation of the following bonds using Lewis dot symbol: (C1, C2) i. Ionic or electrovalent bond ii. Covalent bond iii. Dative or coordinate bond

4.1 LEWIS STRUCTURE

Teaching and learning outcomes

At the end of the lesson, student should be able to

4.1	**Lewis Structure**
d)	Draw Lewis structure of molecules and polyatomic ions with single, double and triple bonds. (C3)
e)	Compare the bond length between single, double and triple bonds. (C2, C4)

4.1 LEWIS STRUCTURE

Teaching and learning outcomes

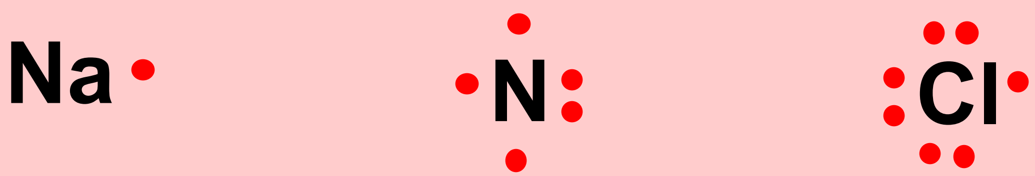
At the end of the lesson, student should be able to

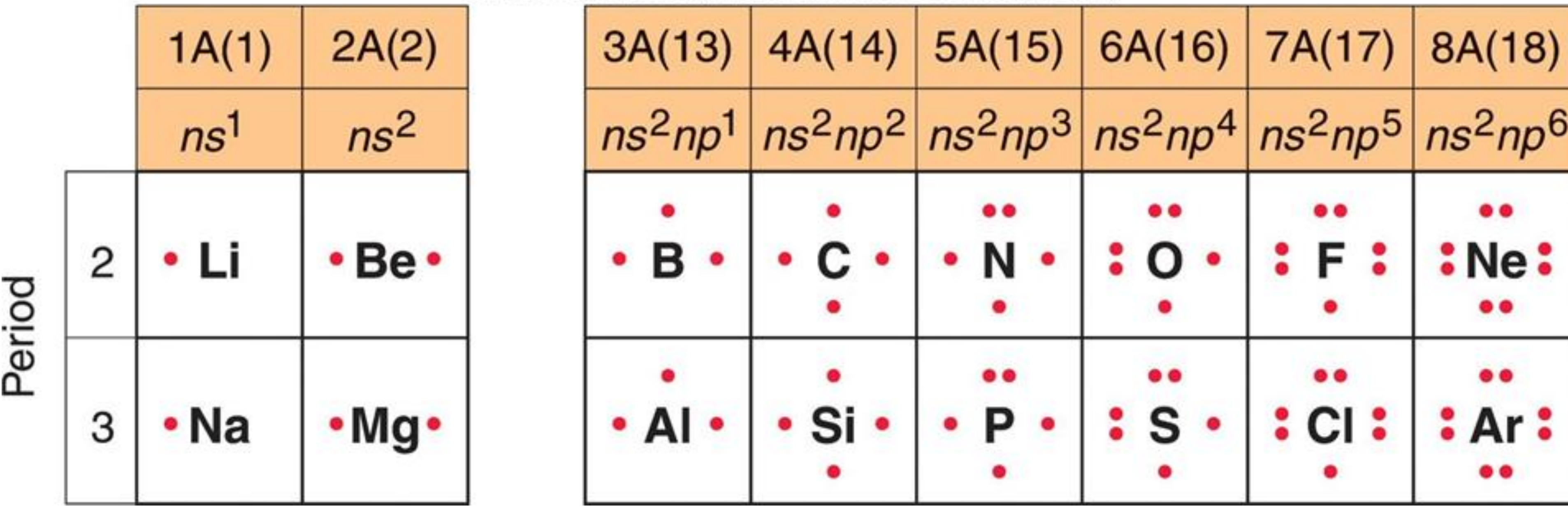
4.1	**Lewis Structure**
f)	Determine the formal charge and the most plausible Lewis structure. (C3)
g)	Explain the exception to the octet rule: (C2, C3) i. Incomplete octet ii. Expanded octet iii. Odd number electron
h)	Illustrate the concept of resonance using appropriate examples. (C2, C3, C4)

LEWIS DOT SYMBOLS

- The valence electrons of main−group elements are represented as dots or cross sign surrounding the symbol of the element

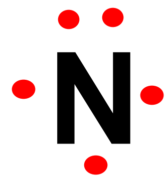
EXAMPLE:



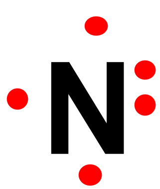

1A(1)
2A(2)
3A(13)
4A(14)
5A(15)
6A(16)
7A(17)
8A(18)
ns1
ns2
ns2np1
ns2np2
ns2np3
ns2np4
ns2np5
ns2np6
Period
2
3
Li
Be
B
C
N
O
F
Ne
Na
Mg
Al
Si
P
S
Cl
Ar

Keep in mind!

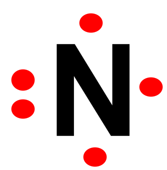
- **Lewis dot–symbol for N:**



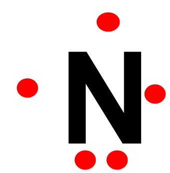
- **Also can be written as:**



or



or



Example 1

4.1

Draw a Lewis electron–dot symbol for each atom.

(a) Rb

(b) Si

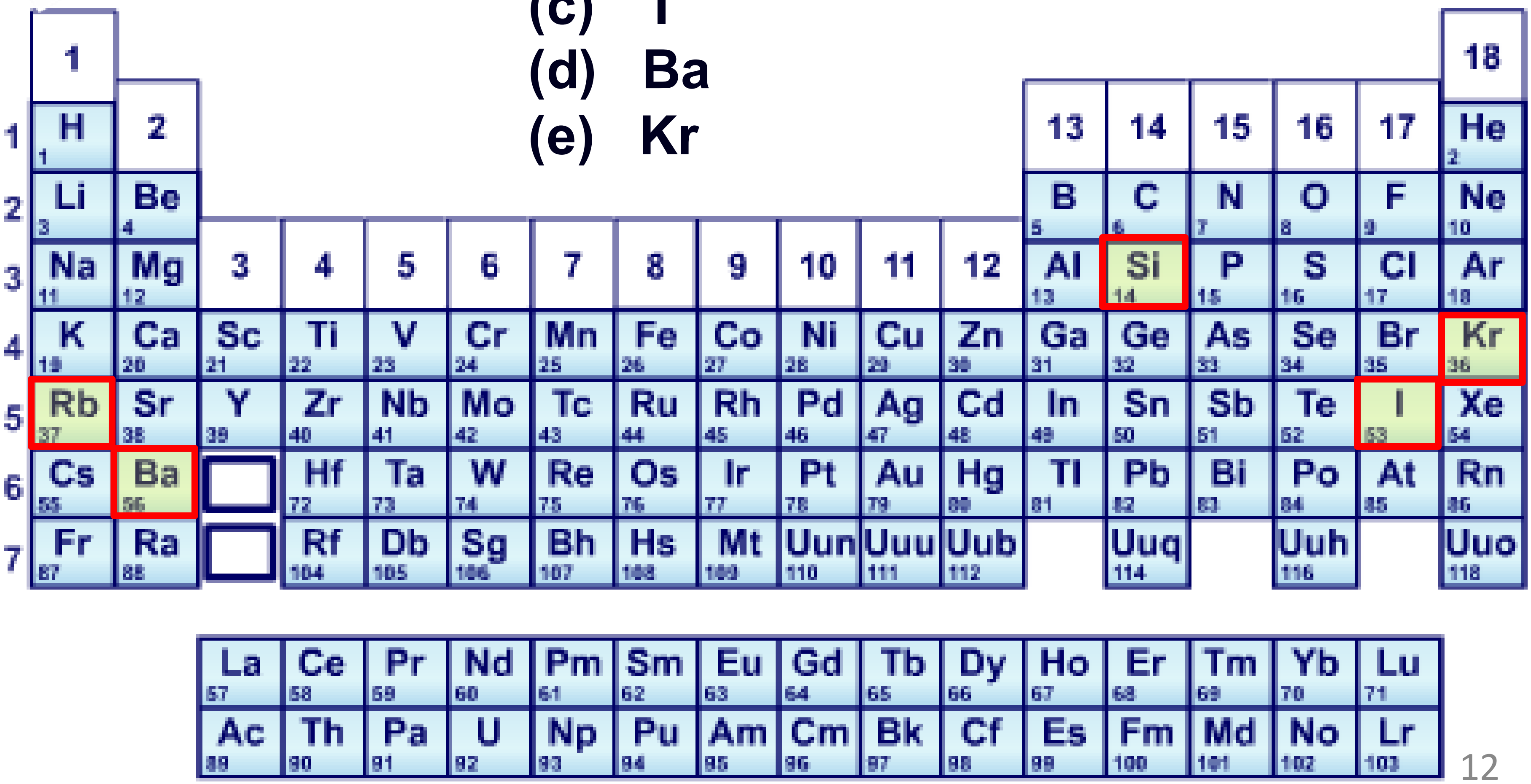
(c) I

(d) Ba

(e) Kr

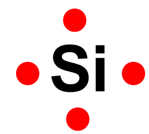
Answer

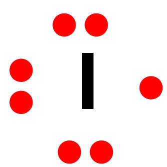
(a) Rb
(b) Si
(c) I
(d) Ba
(e) Kr



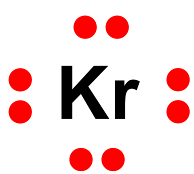
Answer

(a) 

(b) 

(c) 

(d) 

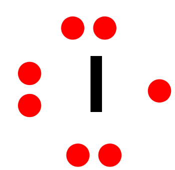
(e) 

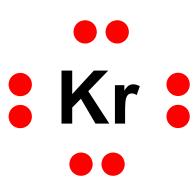
Example 2

4.1

The number unpaired dots provide information about an element bonding behavior:
What information you can get from the following Lewis symbols :

(a) 

(b) 

(c) 

(d) 

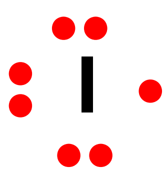
Answer

4.1

(a)

☞ Rb loses one electron to form Rb^+ (charge +1), can form ionic bond.

(b)



☞ I can form one covalent bond.

Answer

4.1

(c) Kr

☞ Kr do not form ionic or covalent bond.

(d) •Ba •

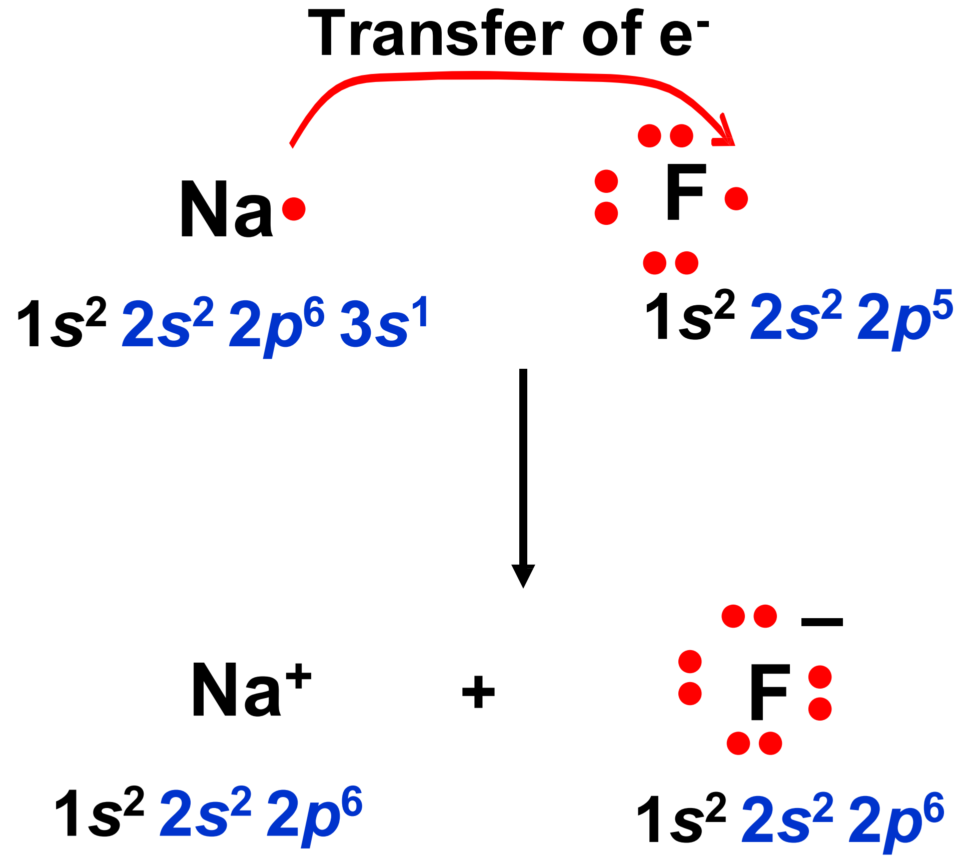
☞ Ba loses two electrons to form Ba^{2+} (charge +2), can form ionic bond.

OCTET RULE

- An atom other than H tends to form bonds (by losing or gaining or sharing e^-) until it is surrounded by eight valence e^-

- The rule works mainly for elements in Period 2

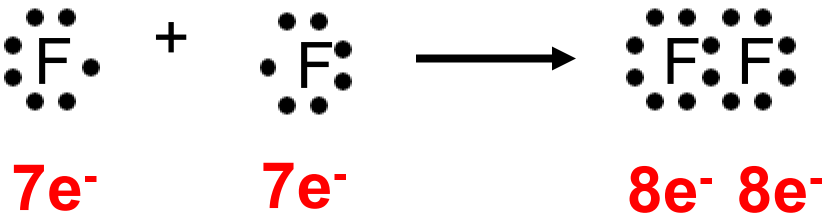
EXAMPLE:

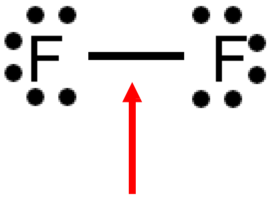


Transfer of e-
Na
F
1s2 2s2 2p6 3s1
1s2 2s2 2p5
Na+
+
F
–
1s2 2s2 2p6
1s2 2s2 2p6

EXAMPLE:

Sharing of e^-





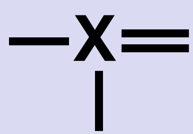
single covalent bond

Example 3

4.1

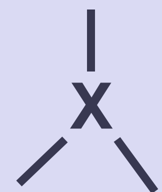
Which of the following bonding patterns does obey the octet rule?

(a)

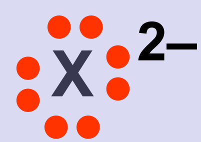


(b)

(c)

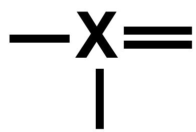


(d)



Answer

4.1

(a) 

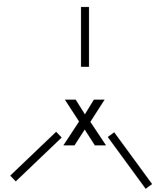
- ☞ X surrounded by 8 electrons
- ☞ Obey octet rule

(b) 

- ☞ X surrounded by 8 electrons
- ☞ Obey octet rule

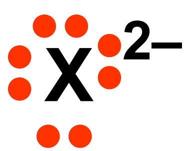
Answer

(c)



- X surrounded by 6 electrons
- **Do not obey octet rule**

(d)



- X surrounded by 8 electrons
- **Obey octet rule**

TYPE OF STABILITY OF IONS

- **Noble gas configuration**
- **Pseudonoble gas configuration**
- **Half–filled orbitals**

NOBLE GAS CONFIGURATION

- Atoms may **lose** or **gain** enough e^- so as to forms **stable** ion with **octet** (or **duplet**) configuration

- The ions formed are stable due to the noble gas configuration
  - ☞ $ns^2\ np^6$

EXAMPLE:

$$Na\cdot \longrightarrow Na^{+} + e^{-}$$

$1s^2\,2s^2\,2p^6\,3s^1$ $\qquad$ $1s^2\,2s^2\,2p^6 = [Ne]$

$$:\ddot{\underset{\cdot\cdot}{Cl}}\cdot + e^{-} \longrightarrow :\ddot{\underset{\cdot\cdot}{Cl}}:^{-}$$

$1s^2\,2s^2\,2p^6\,3s^2\,3p^5$ $\qquad$ $1s^2\,2s^2\,2p^6\,3s^2\,3p^6 = [Ar]$

Example 4

4.1

Write the charge and electron configuration of the ions formed by these elements :

(a) Cl

(b) Na

(c) Ca

(d) N

(e) Br

Answer

4.1

(a) Cl^- : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

(b) Na^+ : $1s^2\ 2s^2\ 2p^6$

(c) Ca^{2+} : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

(d) N^{3-} : $1s^2\ 2s^2\ 2p^6$

(e) Br^- : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

PSEUDONOBLE GAS CONFIGURATION

- **Ions that have stable electronic configurations in which all their orbitals are completely filled with electrons.**

- **The valence electron configuration is** $ns^2np^6nd^{10}$

- **But, electronic configuration is not that of any noble gas**

  ☞ ***pseudonoble gas configuration***

PSEUDONOBLE GAS CONFIGURATION

EXAMPLE:

$$Zn \longrightarrow Zn^{2+} + 2e^{-}$$

Electronic configuration of Zn:

$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2$$

Electronic configuration of Zn^{2+}:

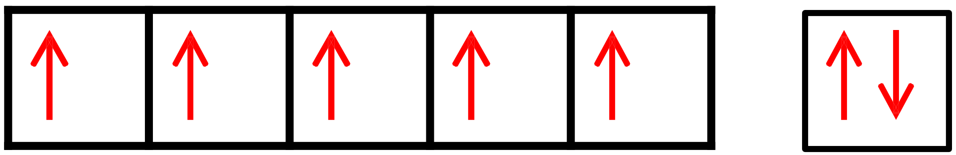
$$1s^2\ 2s^2\ 2p^6\ \color{red}{3s^2\ 3p^6\ 3d^{10}}$$

HALF–FILLED ORBITALS

- Some transition metal atoms form cations that have e^- configuration associated with half–filled *d* orbital (d^5)

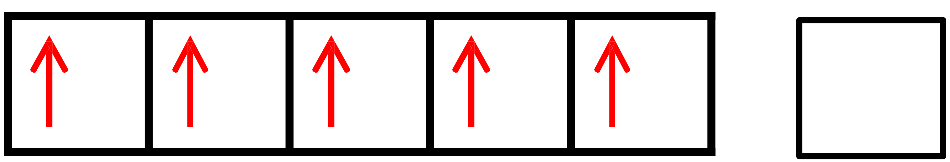
EXAMPLE:

$$Mn \rightarrow Mn^{2+} + 2e^{-}$$



3*d* 4*s*

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2$



3*d* 4*s*

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5$

Example 5

4.1

What type of stability of the electron configuration of ion Fe^{3+}.
Note: Fe (Z = 26)

Answer

4.1

SK01

Fe (Z = 26)

Electron configuration:

$ls^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

Fe^{3+}

Electron configuration:

$ls^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5$

☞ Type of stability = half–filled *d* orbital

CHEMICAL BONDING

- Three major types:

  1. Ionic bond
  2. Covalent bond
  3. Metallic bond

  (will be discussed further in 4.5)

IONIC BOND

- **Attractive electrostatic force between positive and negative ions**

- **Sometimes called electrovalent bond**

Substance formed by Ionic Bonding

Anion / Cation	Cl^-	OH^-	O^{2-}	SO_4^{2-}
Na^+	$NaCl$	$NaOH$	Na_2O	Na_2SO_4
Ca^{2+}	$CaCl_2$	$Ca(OH)_2$	CaO	$CaSO_4$
Al^{3+}	$AlCl_3$	$Al(OH)_3$	Al_2O_3	$Al_2(SO_4)_3$

$AlCl_3$ is a covalent compound
BUT
AlF_3 is an ionic compound

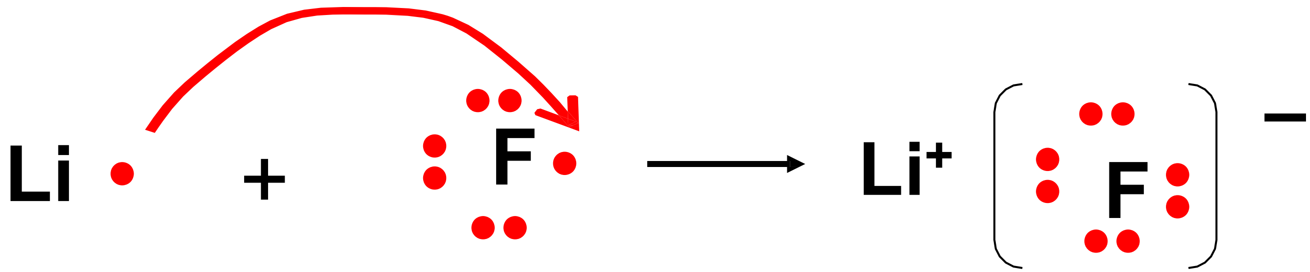
FORMATION OF IONIC BONDS

- Formed when metal combine with nonmetal by electron transfer.

Total number of e^- lost (transferred) by metal atoms	=	Total number of e^- gained by the nonmetal atoms

❑ Lewis electron–dot symbol

EXAMPLE:



- **Lithium atom transfer its valence electron to fluorine and forms lithium ion, Li^+**

- **Fluorine atom accept electron from Li and forms fluorine ion, F^-**

- **The electrostatic forces between Li^+ and F^- forms ionic bond**

Example 6

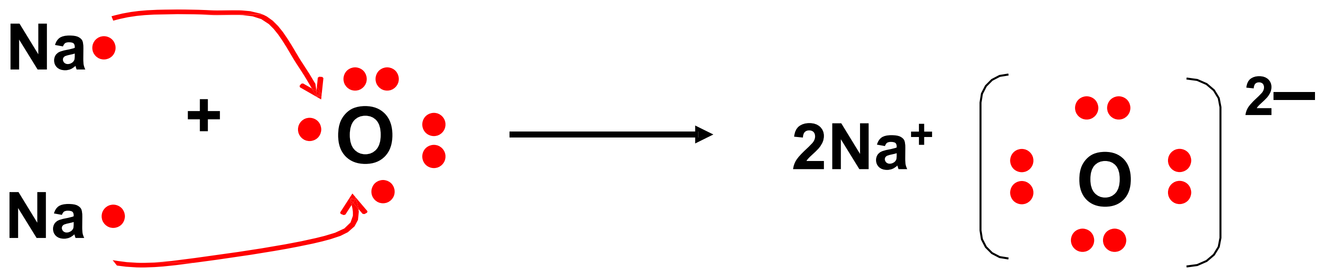
4.1

Use Lewis dot symbols to show the formation of Na_2O.

Answer

4.1

Lewis dot symbols



Electrostatic forces between Na^{+} and O^{2-} forms **ionic bond**

Formula of the compound formed = **Na_2O**

COVAL ENT BOND

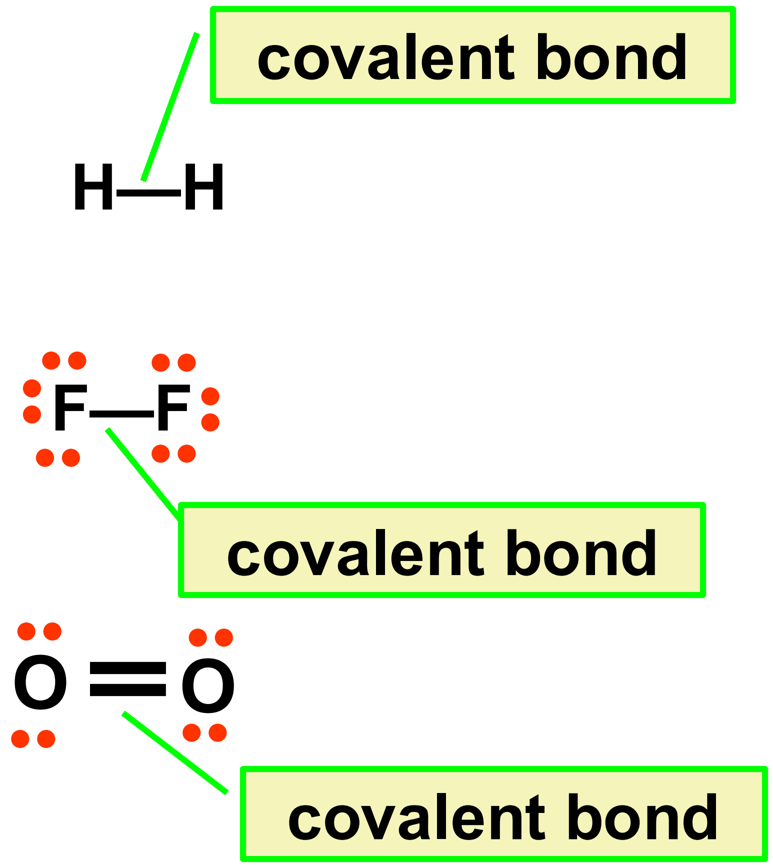
- Formed when nonmetal atoms combine

- Electrons are shared between nonmetal atoms

- The **shared electrons** are counted as **octet** (or duplet) of both atoms and considered to be **localized**

- The electrostatic forces between the shared **electron with the nucleus of both shared atom will form a covalent bonds**

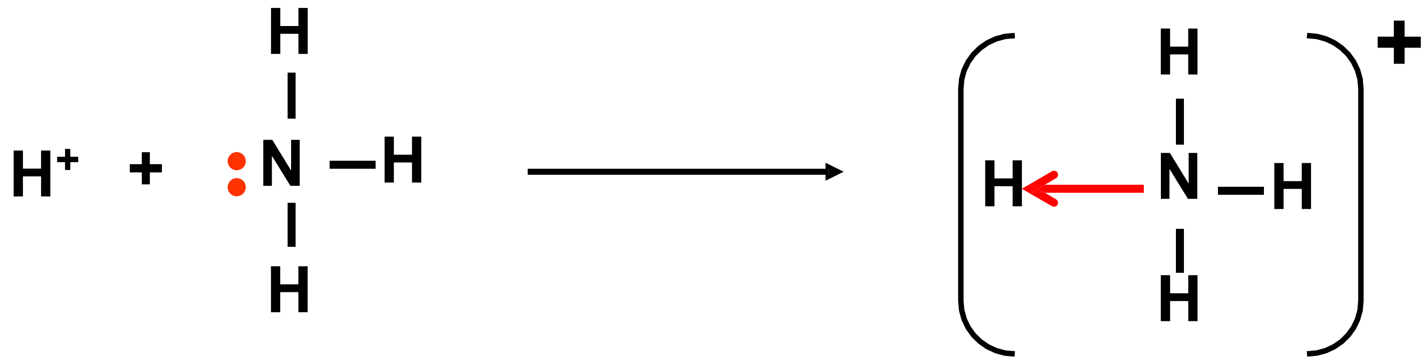
EXAMPLE:



COORDINATE COVALENT BOND (DATIVE BOND)

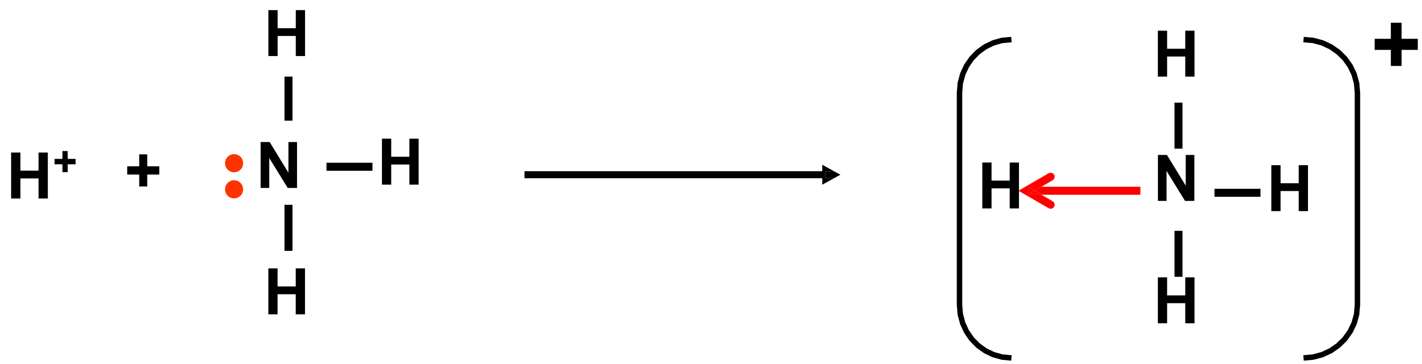
- Formed when one of the atoms **donates both e^-**
- Also called covalent **dative bond**

EXAMPLE:



ammonium ion (NH_4^+)

EXAMPLE: ammonium ion (NH_4^+)



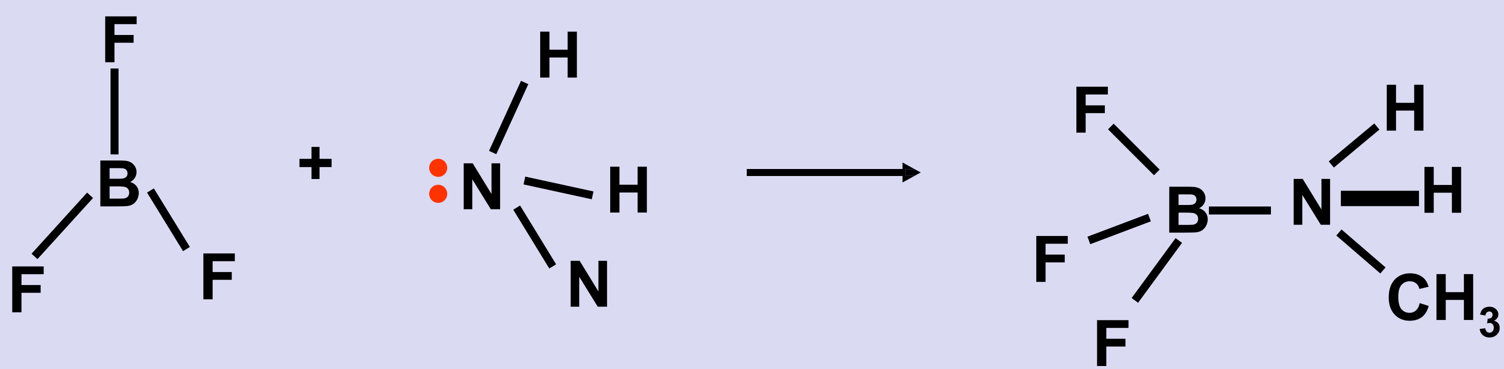
ammonium ion (NH_4^+)

- H^+ has **empty 1s orbital**
- N atom has a **lone pair e^-**
- H^+ accepts an e^- pair from N to form **coordinate covalent bond**

Example 7

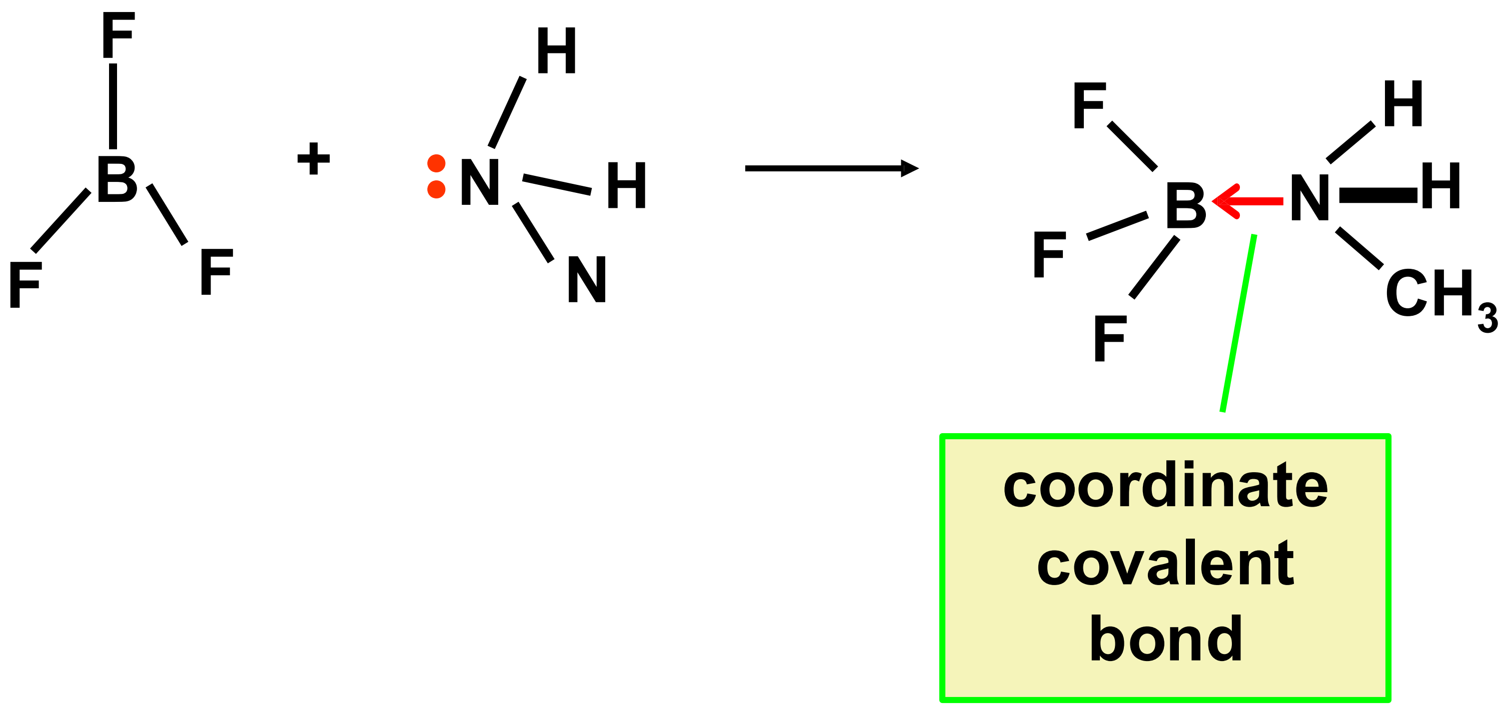
4.1

Boron trifluoride (BF_3) accepts an electron pair from ammonia (NH_3) to form BF_3NH_3.
Show which of the bond is the coordinate covalent bond?



ANSWER

4.1



DRAWING LEWIS STRUCTURE FOR COVALENT MOLECULE

EXAMPLE 1 NF_3

- **Step 1:**
  - **Count the total number of valence e^-**
  - **Add e^- if –ve charge, subtract e^- if +ve charge**

Atom	Number of valence e^-
N x 1	$5e^- \times 1 = 5e^-$
F x 3	$7e^- \times 3 = 21e^-$
	Total = $26e^-$

❑ Step 2:

- Draw skeletal structure of the compound showing the atoms bonded in the molecule
- Put the least electronegative atom in the center

N ☞central atom

F ☞surrounding atom

❑ Step 3:

- **Count the number of electrons which is used to form the covalent bonds and subtract from the total number of electrons.**

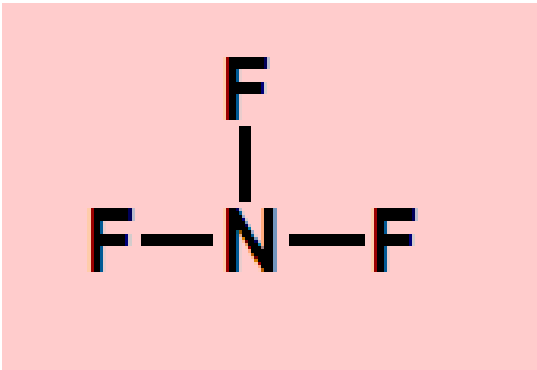
Example :

Number of electrons used to form covalent bond

Total number of valence electrons

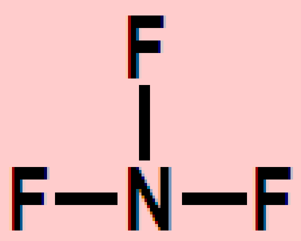
$$26\ e^- - 6\ e^- = 20\ e^-$$

The remaining electrons



❑ Step 4:

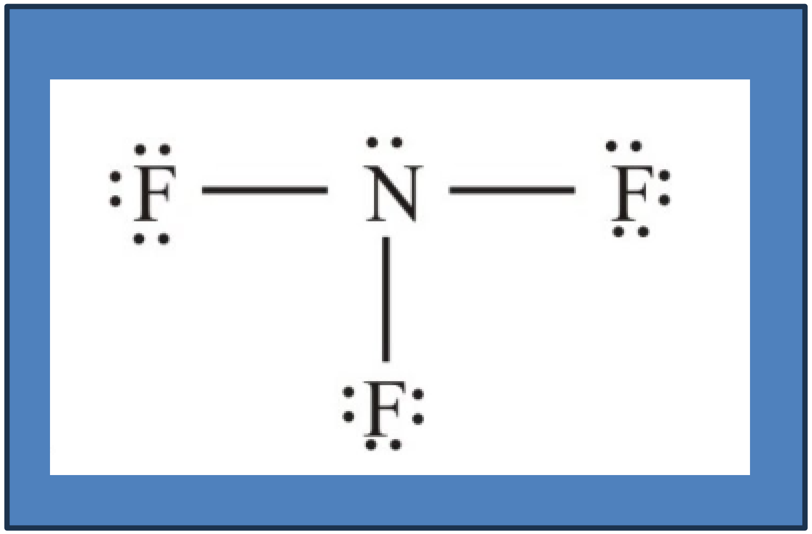
- Complete an **octet** ($8\ e^-$) for **all atoms** except H ($2\ e^-$)
- The surrounding atoms must be octet before placing the remaining electrons at the **central atom**
- **Electrons not involved** in bonding shown as **lone pairs**



Number of electrons at surrounding atoms

$$20\ e^- - 18\ e^- = 2e^-$$

Number of lone pairs electron



Check:

$8e^- \times 3 = 24e^-$

$+ 2e^-$

Total $= 26e^-$

OK!

EXAMPLE 2

NH_4^+

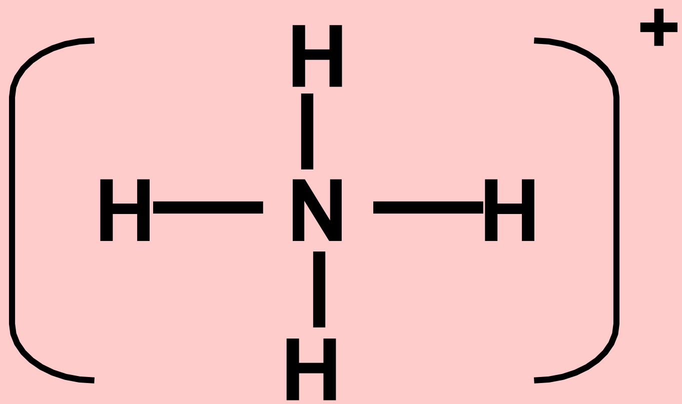
Step 1

- Count the **total number** of **valence e^-**
- Subtract e^- if +ve charge

Atom	Number of valence e^-		
N x 1	$5e^-$ x 1	=	$5e^-$
H x 4	$1e^-$ x 4	=	$4e^-$
+1 charge		=	$-1e^-$
	Total	=	$8e^-$

❑ Step 2:

- Draw the **skeletal structure** of NH_4^+ ion



N ☞ central atom

H ☞ surrounding atom

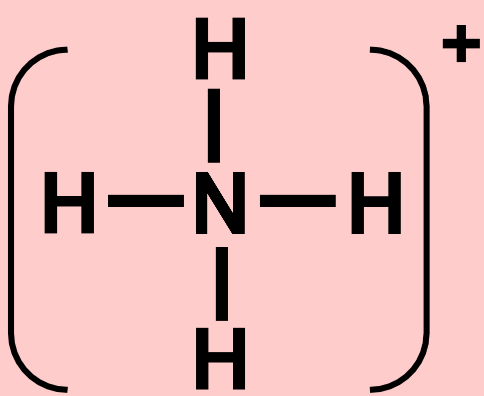
Step 3

- Count the number of electrons which is used to form the covalent bonds and subtract from the total number of electrons.

$$8\ e^- - 8\ e^- = 0\ e^-$$

Step 4

- Complete their duplet for all H atoms ($2\ e^-$)



EXAMPLE 3 CO_3^{2-}

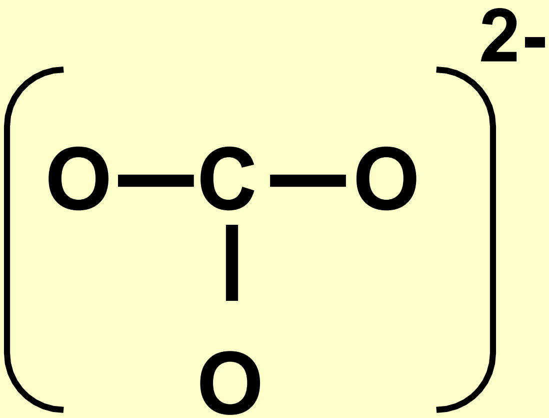
Step 1

- Count **total number** of **valence e^-**
- Add e^- if –ve charge

Atom	Number of valence e^-		
C x 1	$4e^- \times 1$	=	$4e^-$
O x 3	$6e^- \times 3$	=	$18e^-$
-2 charge		=	$+ 2e^-$
	Total	=	$24e^-$

Step 2

- **Draw the skeletal structure of CO_3^{2-} ion**



C	☞	**central atom**
O	☞	**surrounding atom**

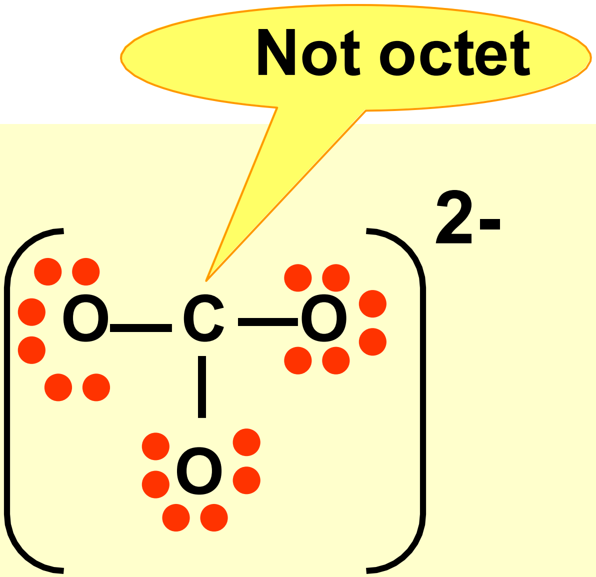
Step 3

- **Count the number of electrons which is used to form the covalent bonds and subtract from the total number of electron.**

$$24\ e^- - 6\ e^- = 18\ e^-$$

Step 4

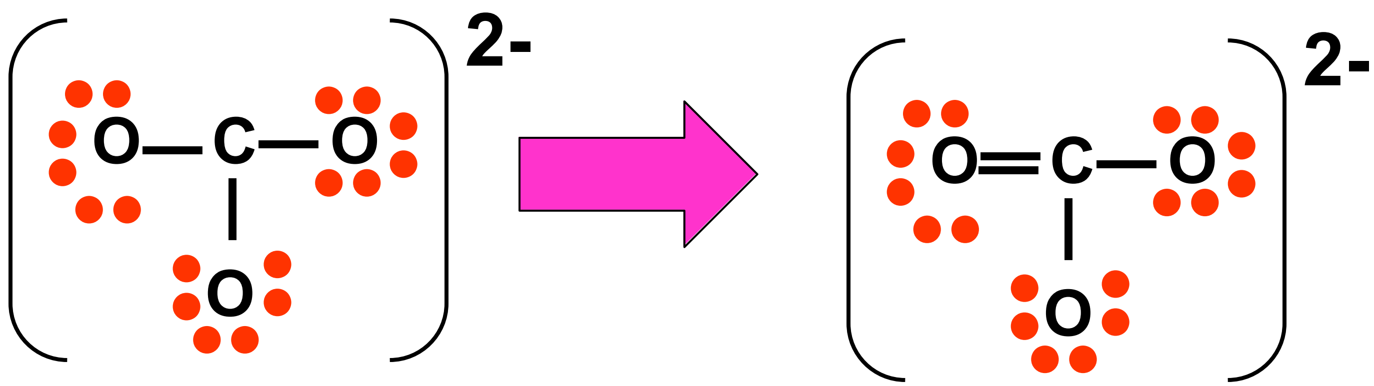
- **Complete an octet for all O atoms (8 e⁻)**



$$18\ e^- - 18\ e^- = 0e^-$$

Step 5

- If the central atom is not **octet yet**, make a **multiple bond** by using a **lone pair** from the **surrounding atoms**



Check:

$8e^- \times 3 = 24e^-$ **OK!**

$\text{Total} = 24e^-$

EXERCISE 1

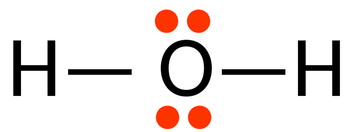
4.1

Draw the Lewis structure for the following compound :

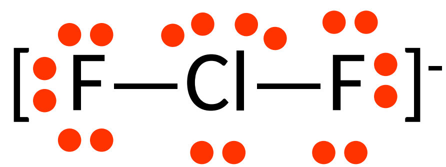
1) H_2O
2) ClF_2^-
3) CN^-

ANSWER

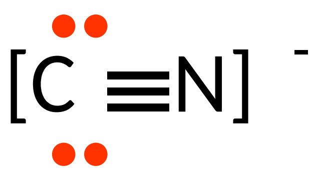
a) H_2O



b) ClF_2^-

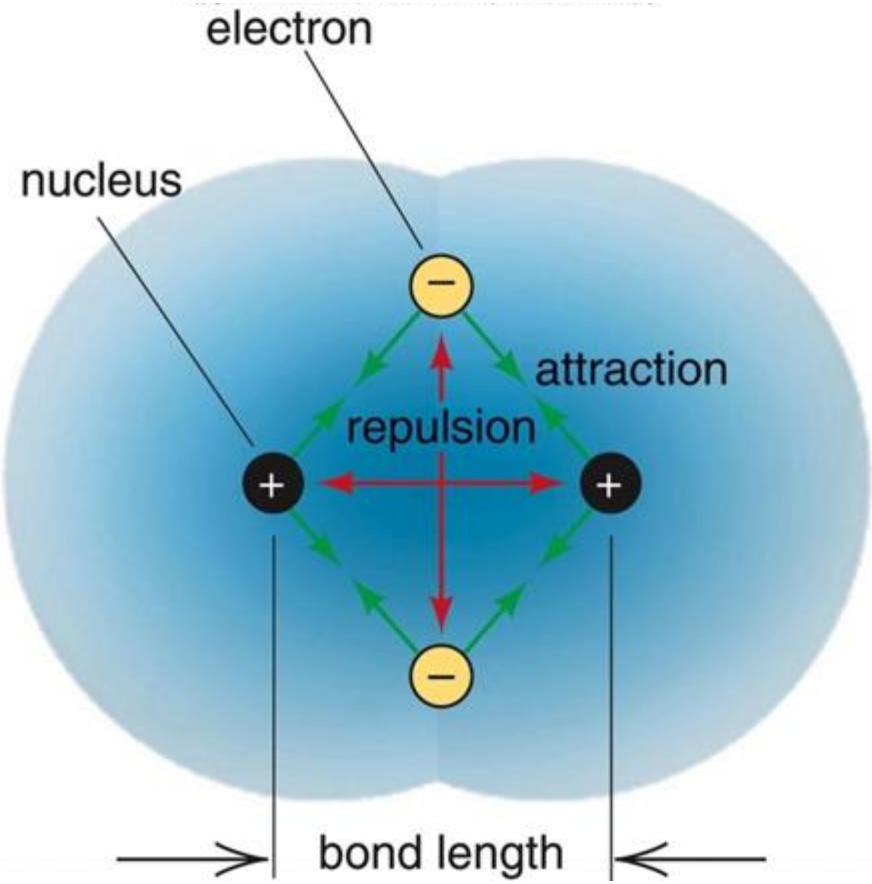


c) CN^-



BOND LENGTH

- **Distance between nuclei of two covalently bonded atoms in a molecule**



- For a given pair of atoms,

Bond length: single > double > triple

EXAMPLE:

Average bond length:

C—C >	C=C >	C≡C
(154 pm)	(134 pm)	(121 pm)

Example 8

4.1

Rank the bonds in each set in order of decreasing bond length:

(a) C=O, C–O, C≡O

(b) N=N, N–N, N≡N

Answer

4.1

(a) C═O, C–O, C≡O

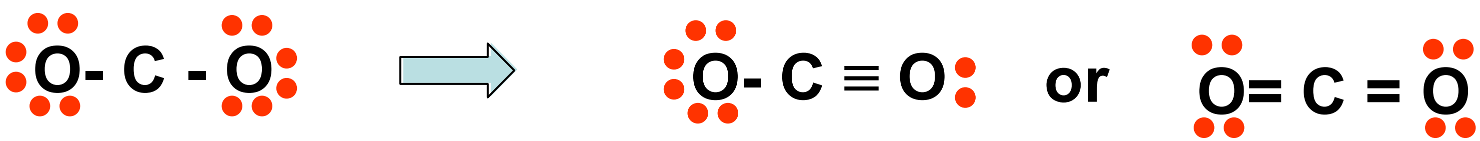
Bond length: C–C > C═C > C≡C

(b) N═N, N–N, N≡N

Bond length: N–N > N═N > N≡N

THE PLAUSIBLE STRUCTURE

- Certain molecules such as CO_2 can have **more** than one lewis structure.



Structure I Structure II

- However, the **most stable structure** is used to represent the molecule which is refered as the **most plausible Lewis structure**

FORMAL CHARGE

- **Difference between the valence e^- in an isolated atom and the number of e^- assigned to that atom in a Lewis structure**

Formal charge of atom =

$$\text{Number of valence } e^- - \left[\text{Number of lone pair } e^- + \tfrac{1}{2} \text{ of bonding } e^- \right]$$

Structure I :

(-1) (0) (+1)

:Ö- C ≡ O:

Formal Charge:

O of C—O ☞[6 – 6 – 1] = -1

O of C≡O ☞[6 – 2 – 3] = +1

C ☞[4 – 0 – 4)] = 0

Note the charge of carbon dioxide

=(−1) + 1 + 0

= 0 ☞CO_2

Structure II :

(0) (0) (0)

O= C = O

Formal Charge:	O of C═O	☞[6 – 4 – 2]	=	0
	O of C═O	☞ [6 – 4 – 2]	=	0
	C	☞[4 – 0 – 4)]	=	0

Note the charge of carbon dioxide

= 0+ 0 + 0

= 0 ☞ CO_2

THE MOST PLAUSIBLE LEWIS STRUCTURE

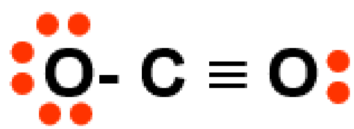
- Select the structure with:

higher priority

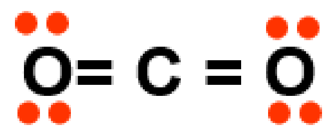
- **Zero** formal charge on all atoms
- **Smaller** formal charge (closest to zero)
- **Negative** formal charges are placed on the **more electronegative** atoms

THE MOST PLAUSIBLE LEWIS STRUCTURE

EXAMPLE 1



or



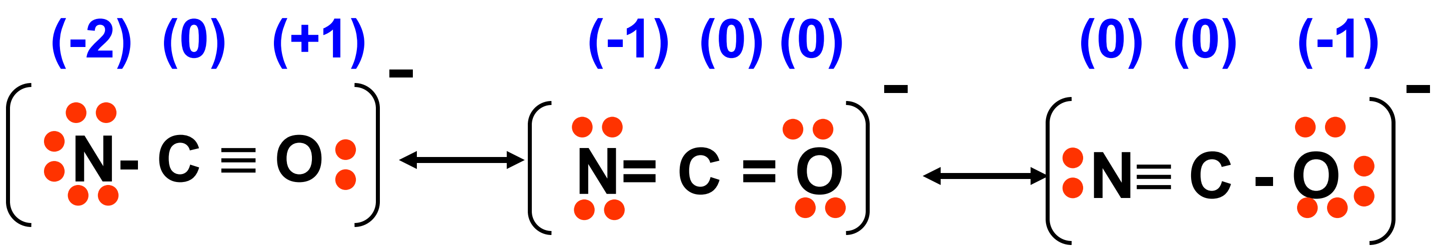
Structure I **Structure II**

The most plausible Lewis structure is II because all atoms have zero formal charge.

EXAMPLE 2

Three possible resonance structure for the ion

☞ **NCO^-**



I II III

III is the most possible resonance structure (stable) because it has fewer formal charges and the -1 formal charge is on the more electronegative atom (O).

EXCEPTION TO OCTET RULE

- Incomplete octet
- Expanded octet
- Odd electron

INCOMPLETE OCTET

- ❑ Occurs when the central atom has less than 8e^- in its valence shell.

- ❑ Elements in group 2 and 13 in period 2 with low metallic properties, do not donate but share the electrons.

- ❑ Be, B and Al cannot achieve octet configuration even after sharing e^- with other atoms.

Example:

BeH_2

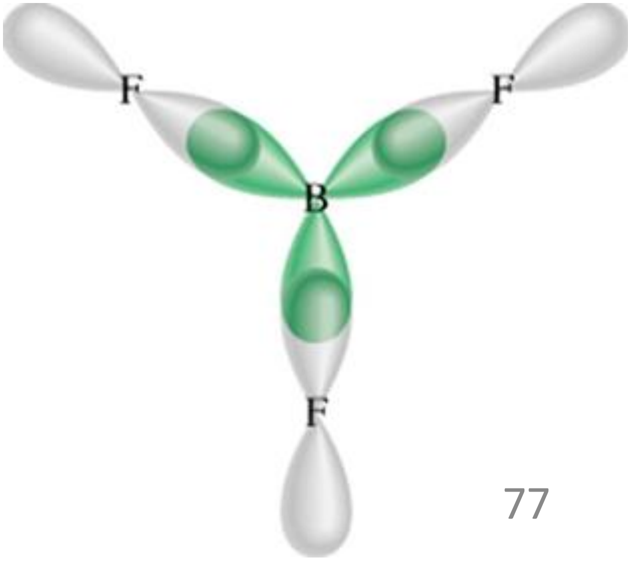
H — Be — H

Be	$2e^-$
2H	$2 \times 1e^-$
	$4e^-$

Example 9

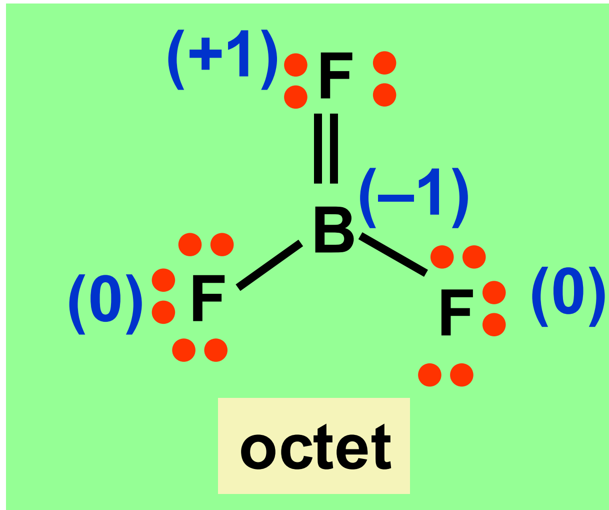
4.1

Draw the Lewis structure of BF_3 that obey the octet rule.
Calculate the formal charge of each atom.
Compare with the incomplete octet Lewis structure of BF_3 .



Answer

4.1

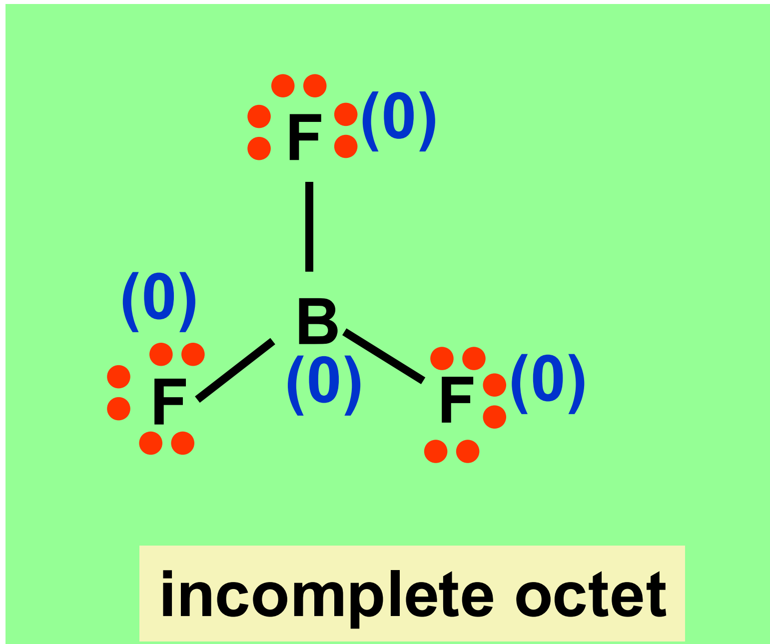


Formal Charge:

F of B—F	☞[7 – 6 – 1)]	=	0
F of B═F	☞[7 – 4 – 2)]	=	+1
B	☞[3 – 0 – 4)]	=	–1

Answer

4.1



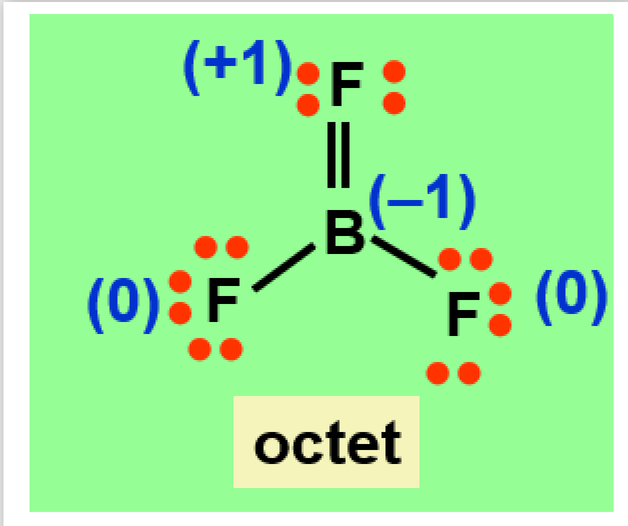
Formal charge:

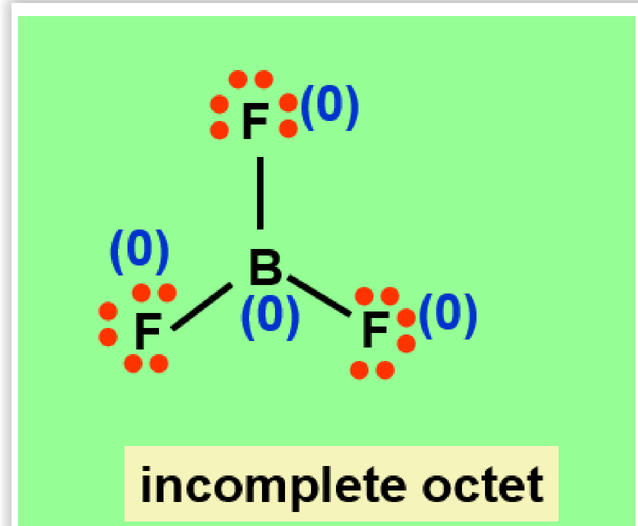
F ☞[7 – 6 – 1)] = 0

B ☞[3 – 0 – 3)] = 0

Answer

4.1





Structure with incomplete octet is more stable because it has zero formal charges at all atoms.

For B=F : formal charge of more electronegative F is +1 while formal charge of less electronegative B is -1.

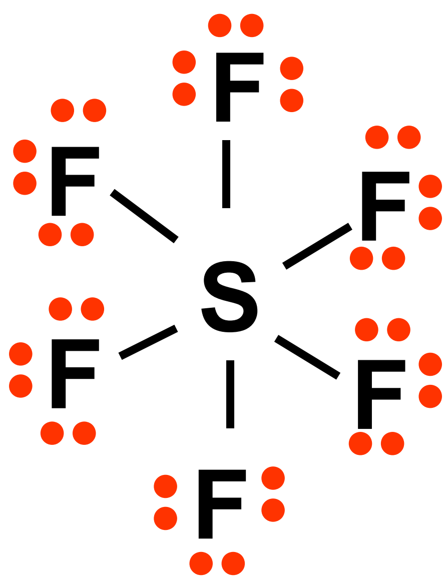
Thus the Lewis structure with complete octet B=F is less stable)

EXPANDED OCTET

- Occurs when central atom have **more than eight e^-** in its valence shell.
- Usually involves non-metal atoms of 3rd period and beyond which have empty 3d subshell.

EXAMPLE:

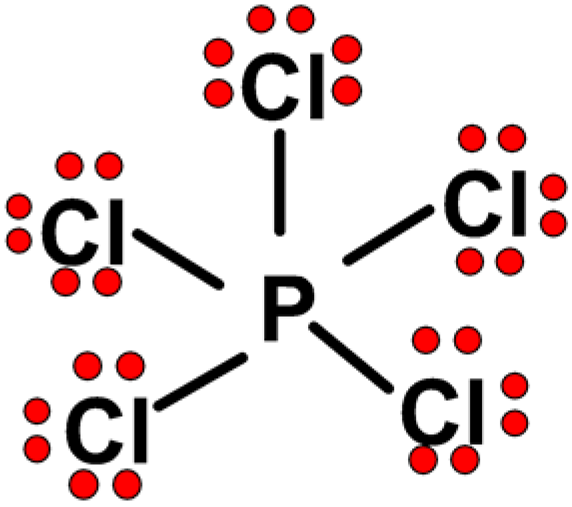
SF_6



$$\begin{array}{lcl} S & = & 6\ e^- \\ 6F & = & 6 \times 7e^- \\ \hline & & 48e^- \end{array}$$

EXAMPLE:

PCl_5



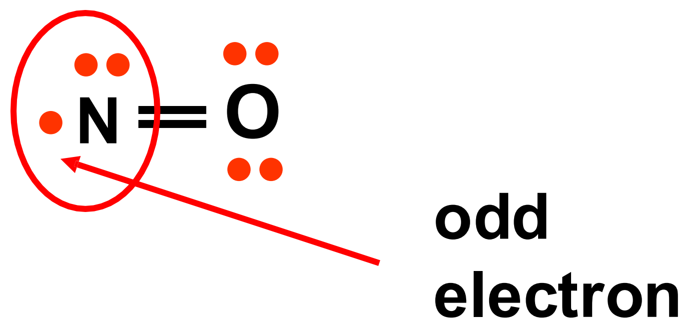
P	$= 5e^-$
5Cl	$= 5x7e^-$
	$40e^-$

ODD–ELECTRON MOLECULES

- Contain an unpaired e^-

EXAMPLE:

NO

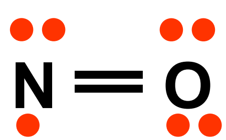


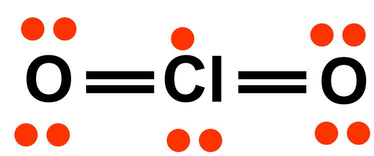
N	$5e^-$
O	$6e^-$
	$11e^-$

Keep in mind!

Most odd–electron molecule have a central atom from an odd–numbered group, such as N (Group 15) and Cl (Group 17)

EXAMPLE:

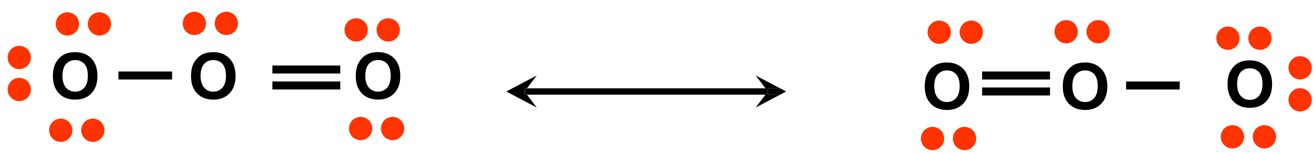




RESONANCE STRUCTURE

- **Two or more Lewis structures for a single molecule that cannot be represented accurately by only one Lewis structure**

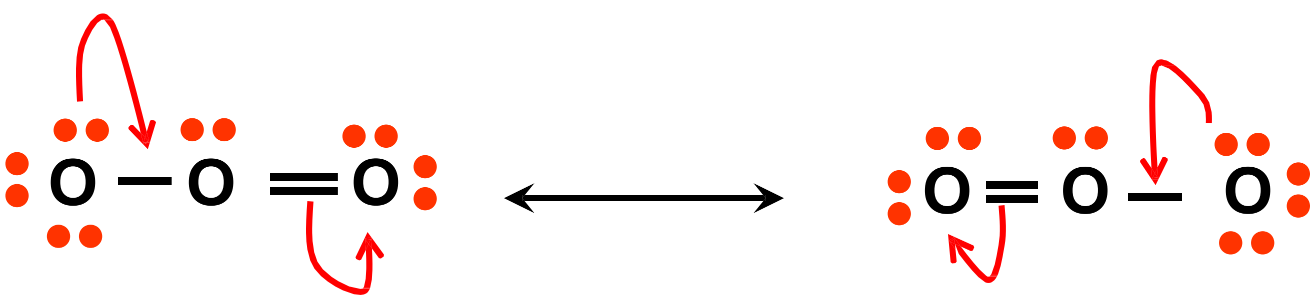
EXAMPLE: **ozone (O_3)**



I II

- **The structures have same relative placement of atoms but different locations of bonding and lone e^- pairs**

EXAMPLE: **ozone (O_3)**

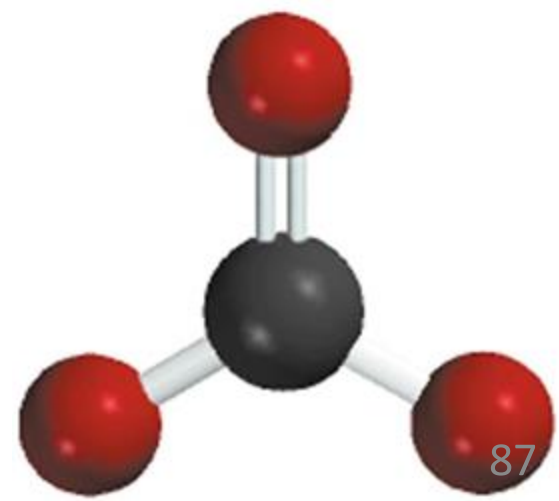


I **II**

Example 10

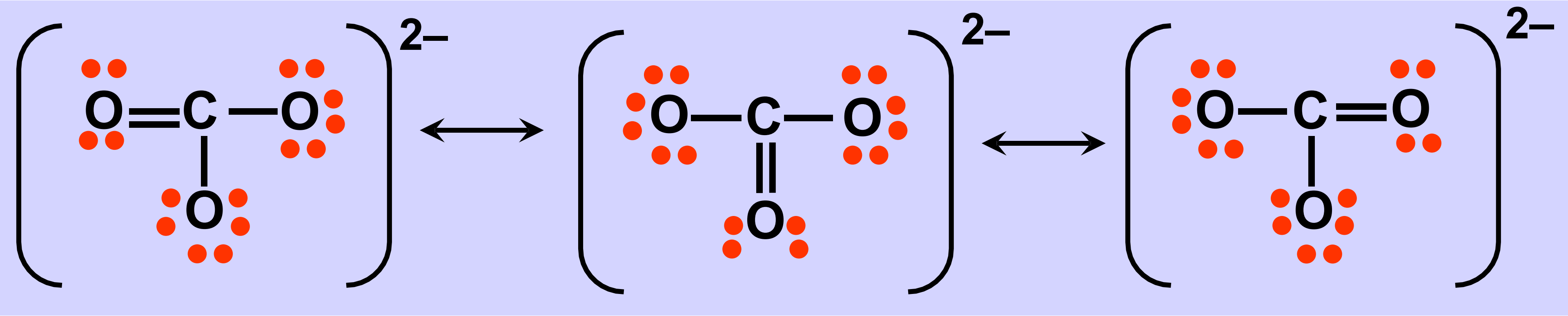
4.1

What are the resonance structures of the carbonate (CO_3^{2-}) ion?



Answer

4.1



CHAPTER 4.2 : OVERVIEW

Molecular Shape & Polarity

- **VSEPR Theory**
- **Basic Molecular Shape**
  - Linear
  - Trigonal Planar
  - Tetrahedral
  - Trigonal bipyramidal
  - Octahedral
- **Molecular Shape**
  - V-shaped
  - Trigonal pyramidal
  - T-Shaped
  - See-saw
  - Square pyramidal
  - Square planar
- **Bond Angle**
- **Dipole Moment**
- **Molecular & Bond Polarity**
  - Dipole moment (μ)
  - Resultant Dipole moment (μ)

4.2 MOLECULAR SHAPE AND POLARITY

Teaching and learning outcomes

At the end of the lesson, student should be able to

4.2	**Molecular shape and polarity**
a)	Explain Valence Shell Electron Pair Repulsion theory (VSEPR) (C2, C3)
b)	Draw the basic molecular shapes (C1) i. Linear ii. Trigonal planar iii. Tetrahedral iv. Trigonal bipyramidal v. Octahedral
c)	Predict the shapes of molecule and bond angles in a given species. (C2, C3)
d)	Explain bond polarity and dipole moment. (C2, C3)
e)	Deduce the polarity of molecules based on the shapes and the resultant dipole moment. (C4)

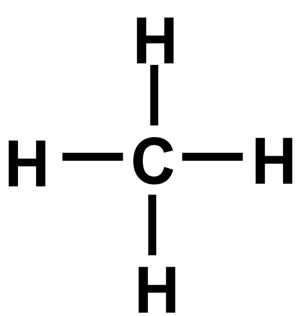
VSEPR THEORY

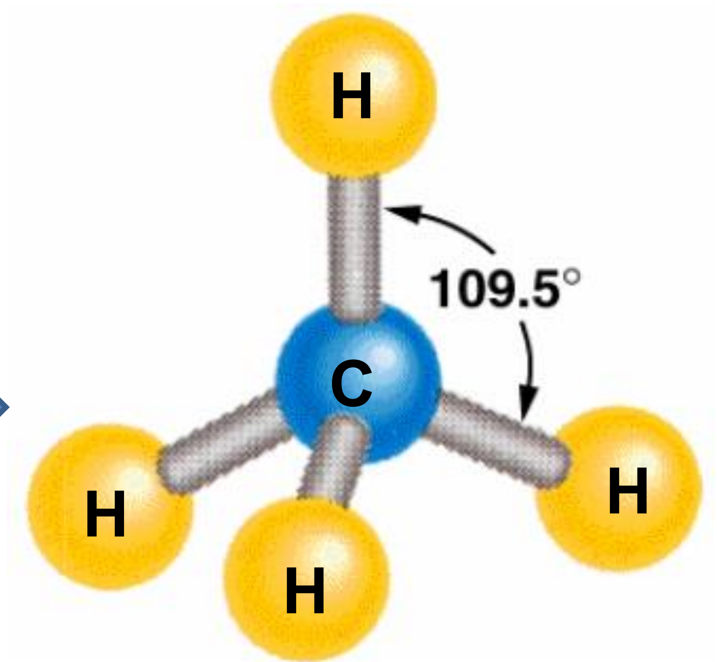
- **Valence–Shell Electron–Pair Repulsion Theory**

- **Each group of valence electrons around a central atom is located as far away as possible from the others in order to minimize repulsion**

- The theory is used to **predict** the **molecular shape** from the Lewis structure

EXAMPLE:

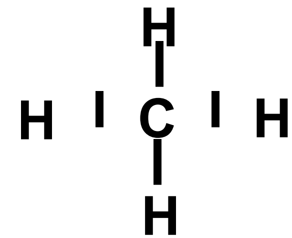
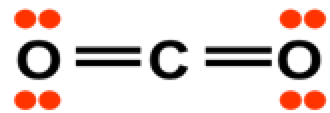
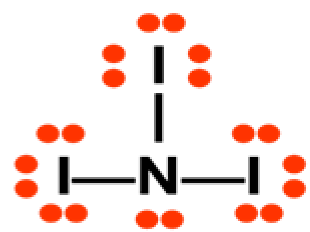
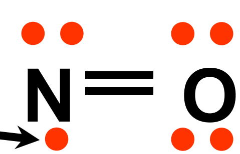




Lewis structure

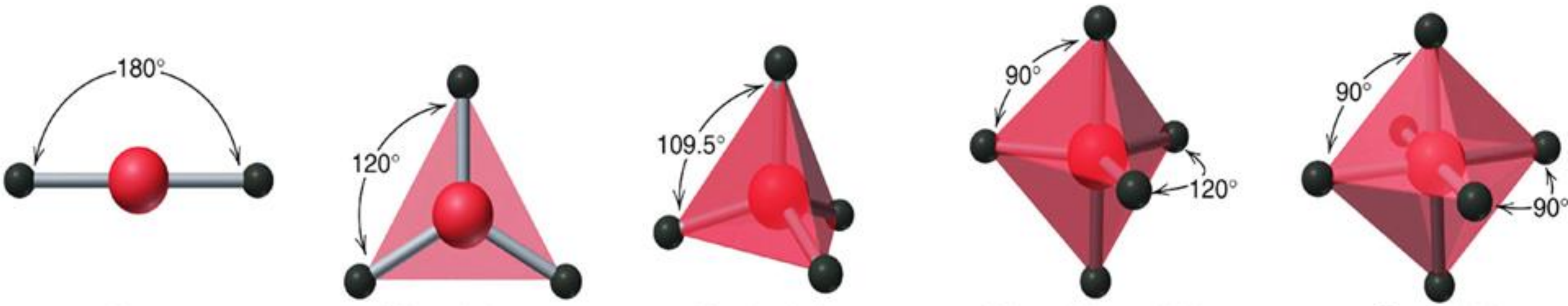
molecular shape

The followings count as one e$^-$ group:

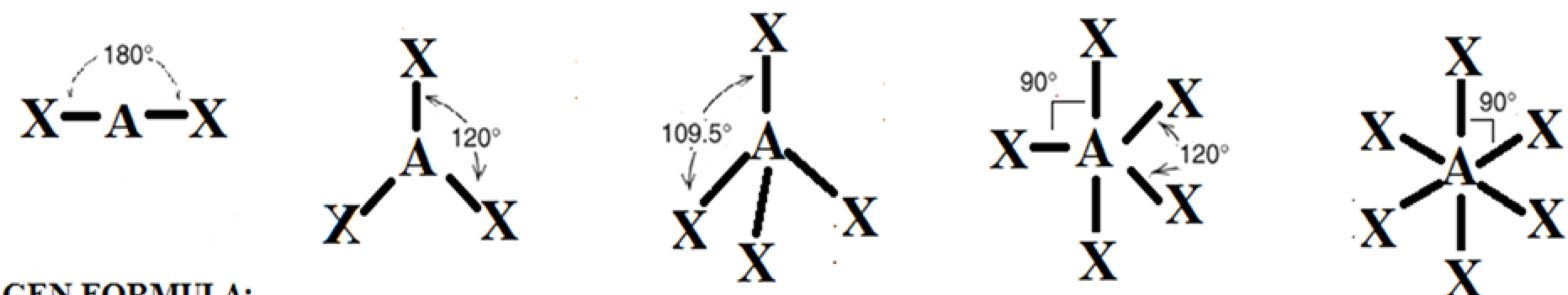
❶ **Bonding pair** **- a single bond** **- a double bond** **- a triple bond**		☞ e$^-$ groups = 4
		☞ e$^-$ groups = 2
❷ **lone pair**		☞ e$^-$ groups = 4
		☞ e$^-$ groups = 4
❸ **lone** e^-		☞ e$^-$ groups = 3

FIVE BASIC SHAPES

❶ Linear ❷ Trigonal planar ❸ Tetrahedral ❹ Trigonal bipyramidal ❺ Octahedral



Linear | Trigonal planar | Tetrahedral | Trigonal bipyramidal | Octahedral



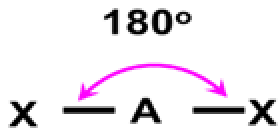
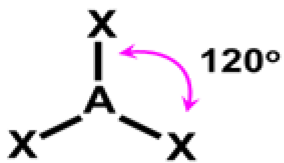
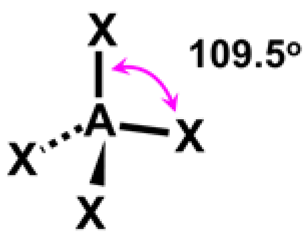
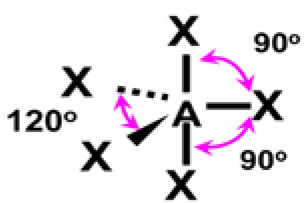
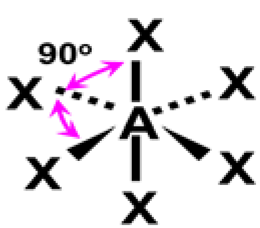
GEN FORMULA:

AX_2 AX_3 AX_4 AX_5 AX_6

ELECTRON GROUP ARRANGEMENT

- ❑ **Determined by the number of e^- groups around the central atom**

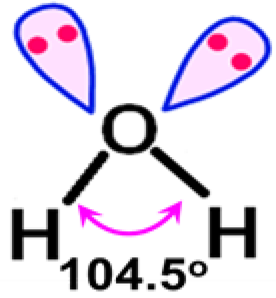
ELECTRON GROUP ARRANGEMENT

Geometry	e^- Groups	Arrangement
	**2**	**linear**
	3	**trigonal planar**
	4	**tetrahedral**
	5	**Trigonal bipyramidal**
	6	**octahedral**

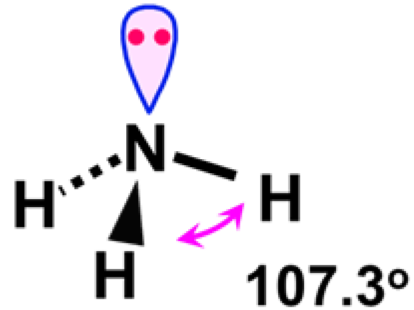
ELECTRON GROUP REPULSION

- **Order:**

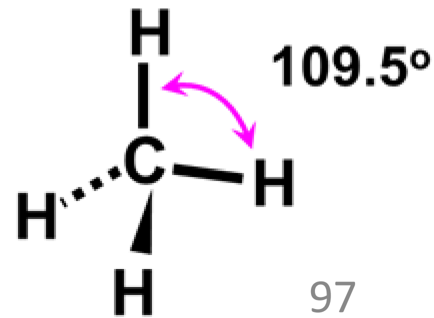
lone pair – lone pair >



lone pair – bonding pair >



bonding pair – bonding pair



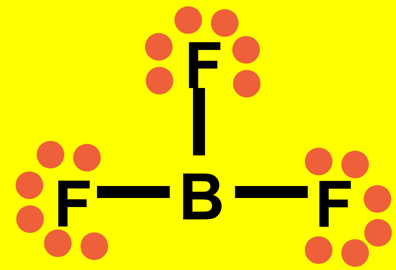
☞ **EXAMPLE: BF_3**

❶ LEWIS STRUCTURE

❷ ELECTRON GROUP ARRANGEMENT

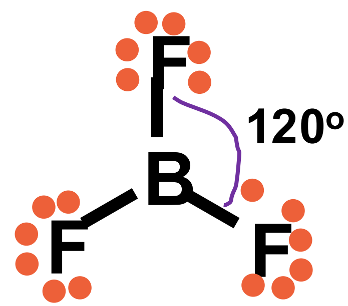
❸ MOLECULAR SHAPE

Lewis structure



☞ **Electron group = 3**

e- group arrangement = trigonal planar

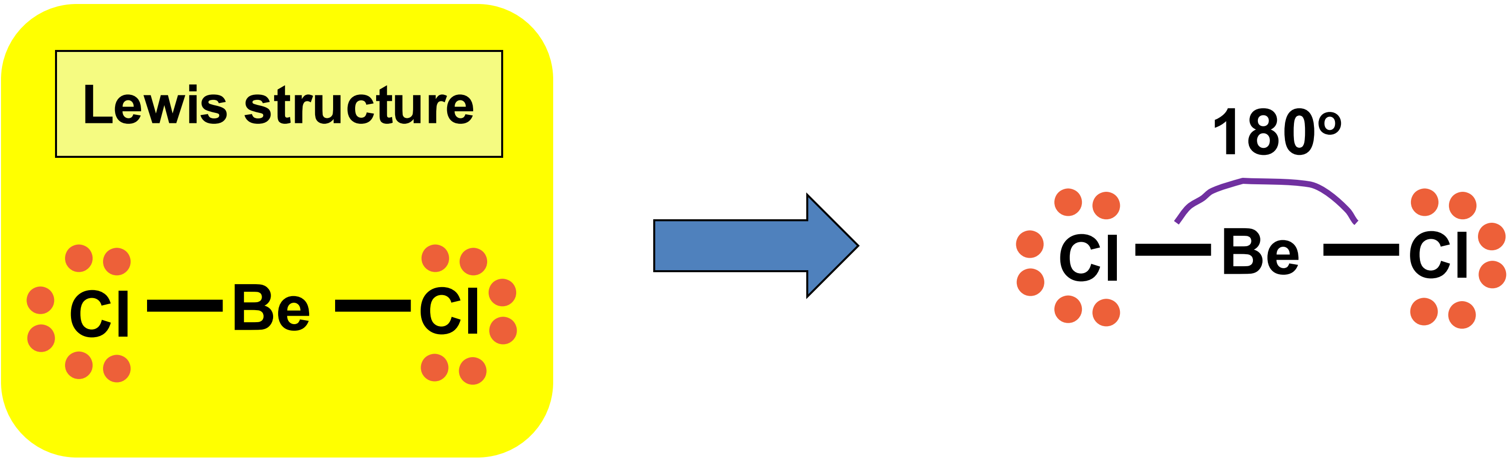


☞ **trigonal planar**

bond angle: 120°

TWO ELECTRON GROUPS

EXAMPLE: $BeCl_2$ (gaseous beryllium chloride)

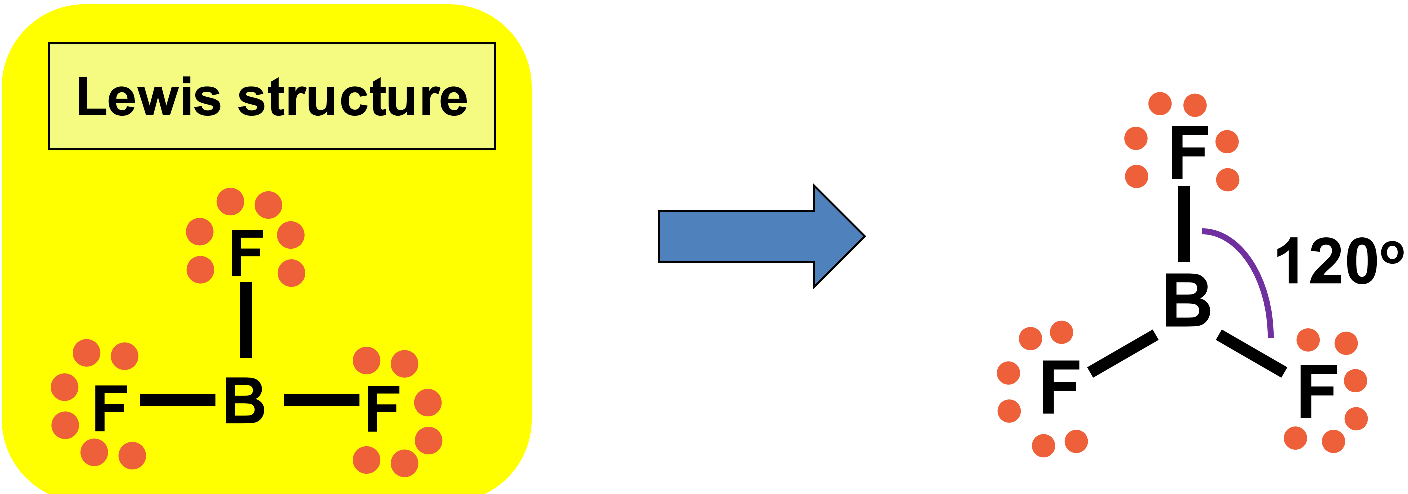


- ❑ **e^- groups arrangement ☞ linear**
- ❑ **Molecular shape ☞ linear**

bond angle: 180°

THREE ELECTRON GROUPS

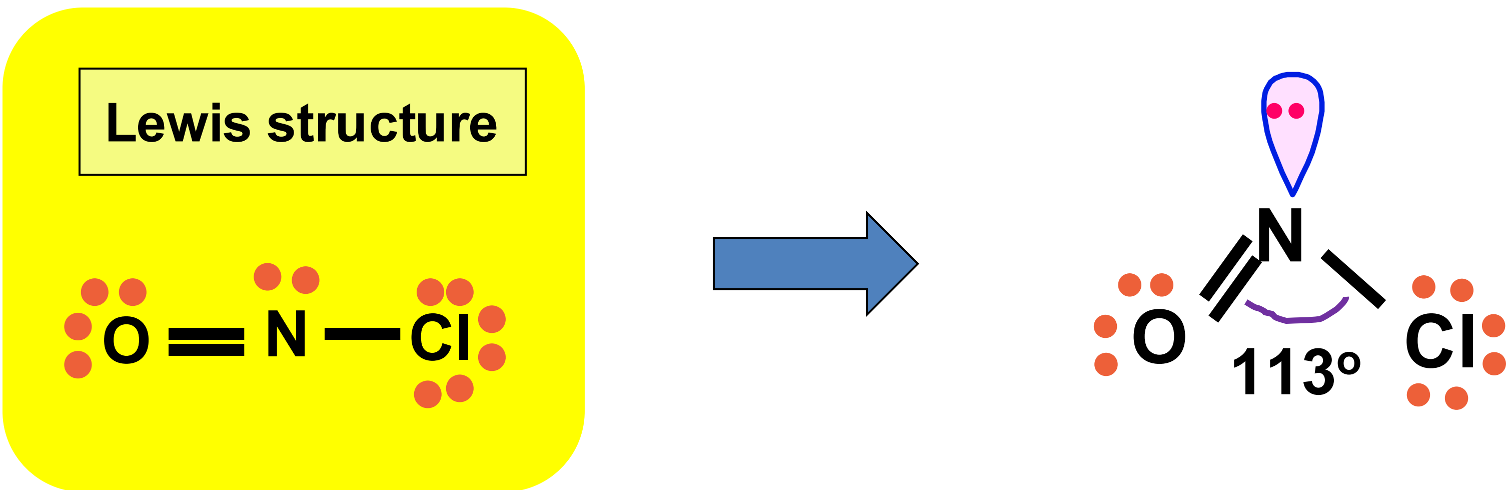
EXAMPLE: BF_3 (boron trifluoride)



- e^- groups arrangement ☞ trigonal planar
- repulsion of all **bonding pair - bonding pair** electrons are equal.
- Molecular shape ☞ trigonal planar

  bond angle: 120°

EXAMPLE: NOCl (nitrosyl chloride)

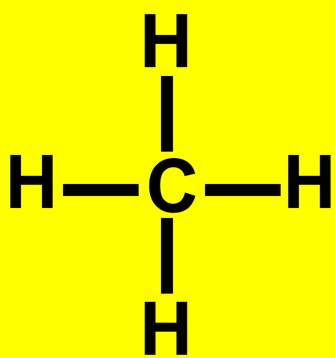


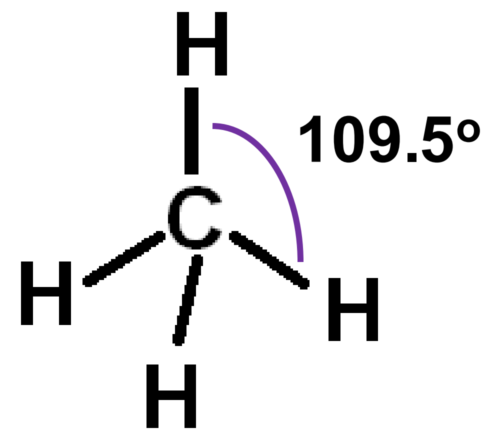
- e^- groups arrangement ☞ trigonal planar
- repulsion of lone pair-bonding pair > bonding pair - bonding pair.
- Molecular shape ☞ V–shaped (bent)
  bond angle: < 120°

FOUR ELECTRON GROUPS

EXAMPLE: CH_4 (methane)

Lewis structure

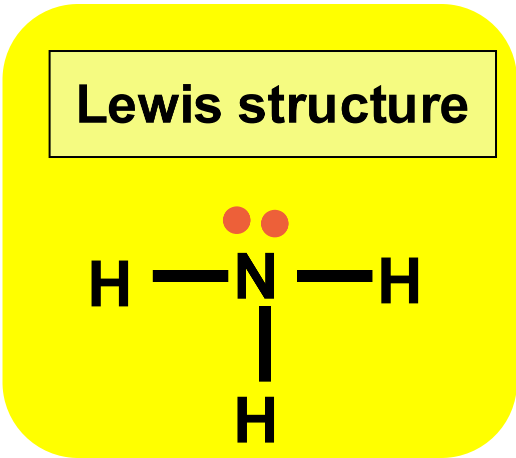


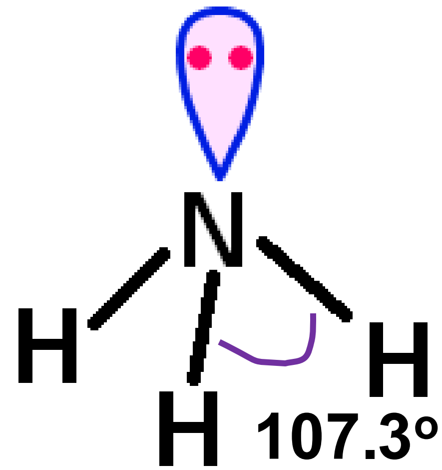


- ❑ **e⁻ groups arrangement ☞ tetrahedral**
- ❑ **Molecular shape ☞ tetrahedral**

bond angle: 109.5°

EXAMPLE: NH_3 (ammonia)

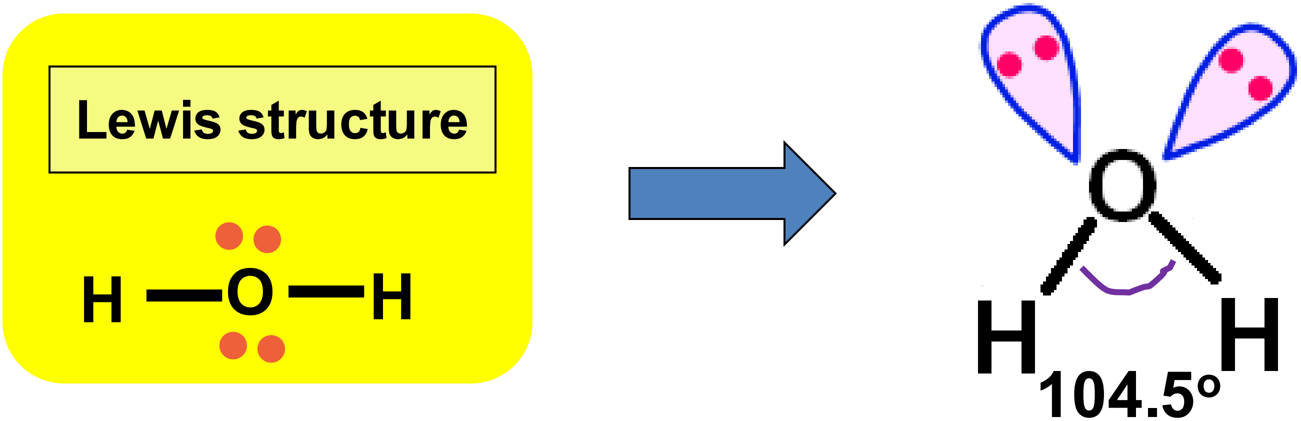




- e⁻ groups arrangement ☞ tetrahedral
- repulsion of lone pair-bonding pair > bonding pair - bonding pair
- Molecular shape ☞ trigonal pyramidal

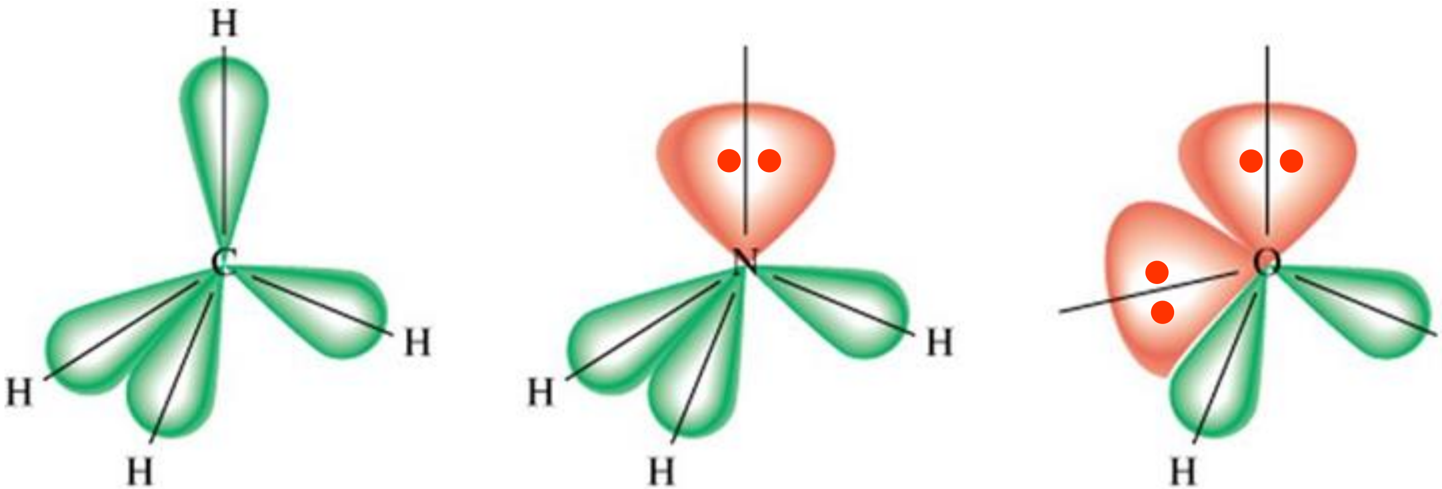
  bond angle: < 109.5° (107.3°)

EXAMPLE: H_2O (water)

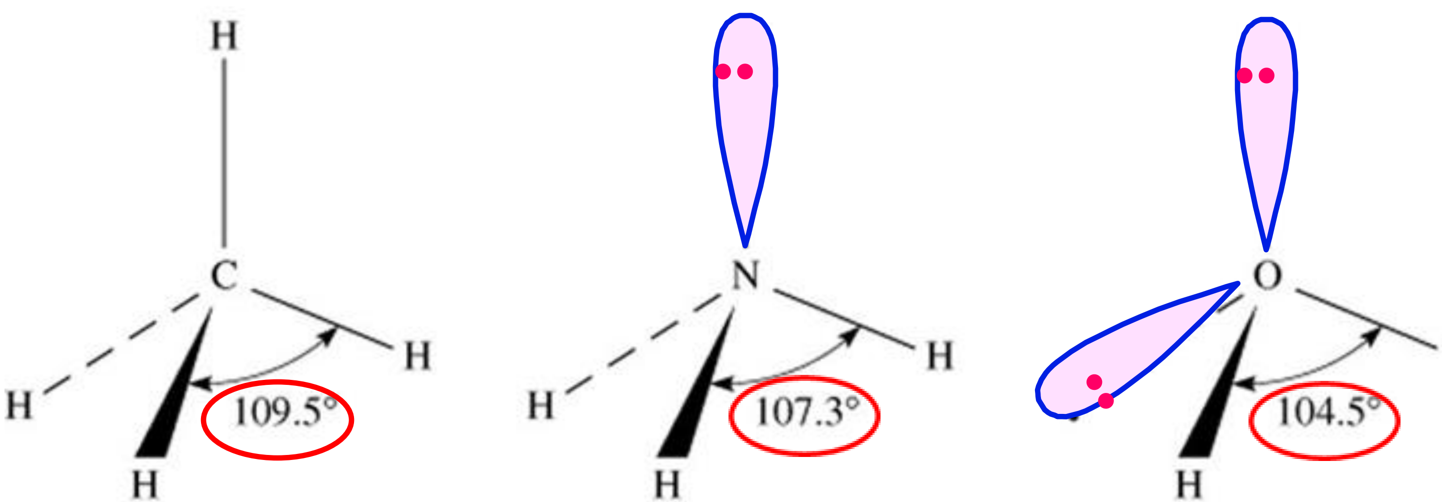


- ❑ e^- groups arrangement ☞ **tetrahedral**
- ❑ repulsion of **lone pair-lone pair** > **lone pair-bonding pair** > **bonding pair-bonding pair**
- ❑ Molecular shape ☞ **V–shape** (**bent**)
  bond angle: **< 109.5°** (104.5°)

TETRAHEDRAL ELECTRON GROUP ARRANGEMENT



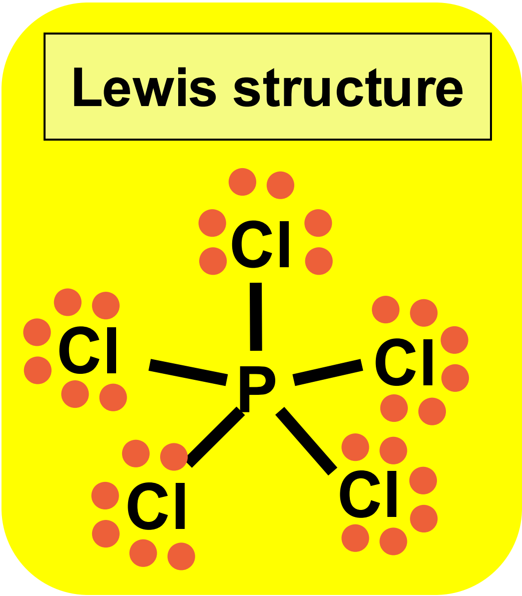
tetrahedral | trigonal pyramidal | V shaped

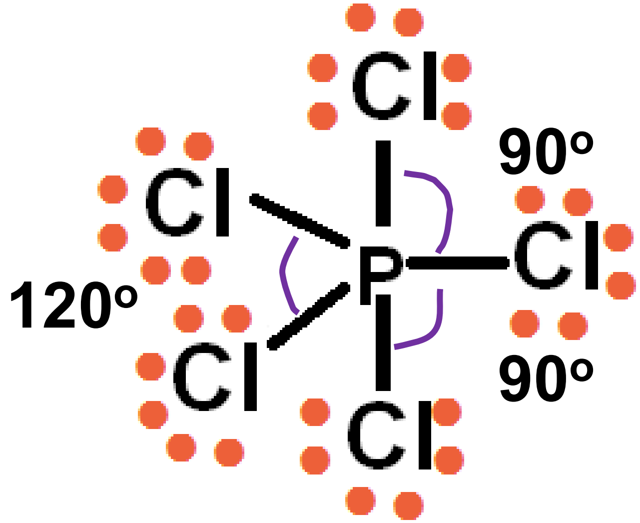


bonding-pair vs. bonding pair repulsion < **lone-pair vs. bonding pair repulsion** < **lone-pair vs. lone pair repulsion**

FIVE ELECTRON GROUPS

EXAMPLE: PCl_5 (phosphorus pentachloride)

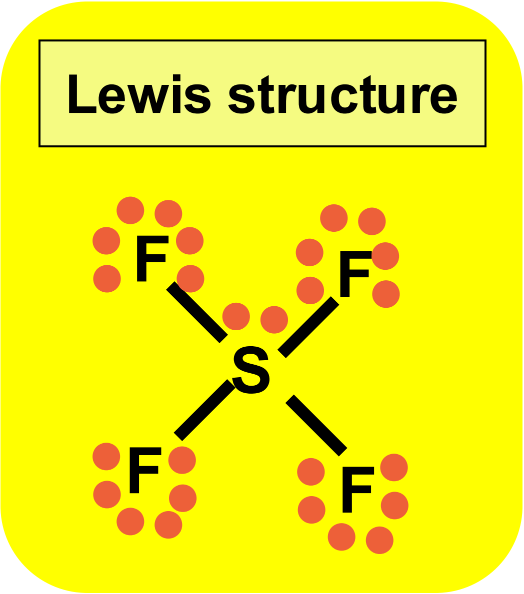


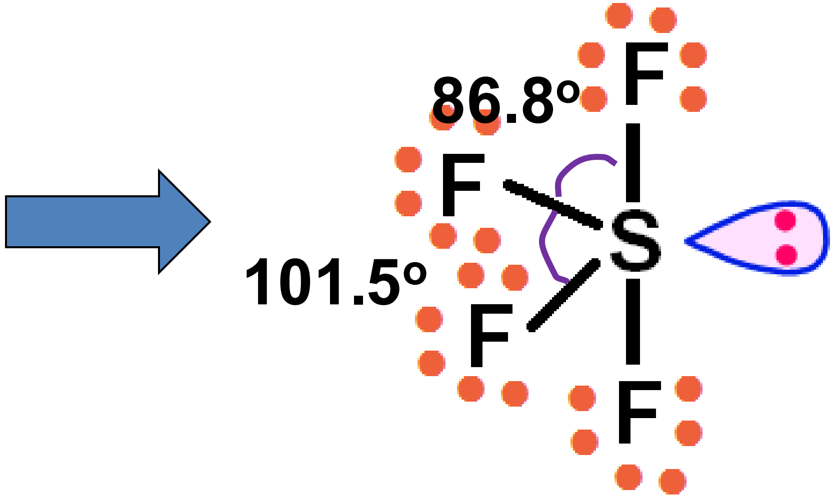


- ❑ e⁻ groups arrangement ☞ **trigonal bipyramidal**
- ❑ Molecular shape ☞ **trigonal bipyramidal**

bond angle: 120°, 90°

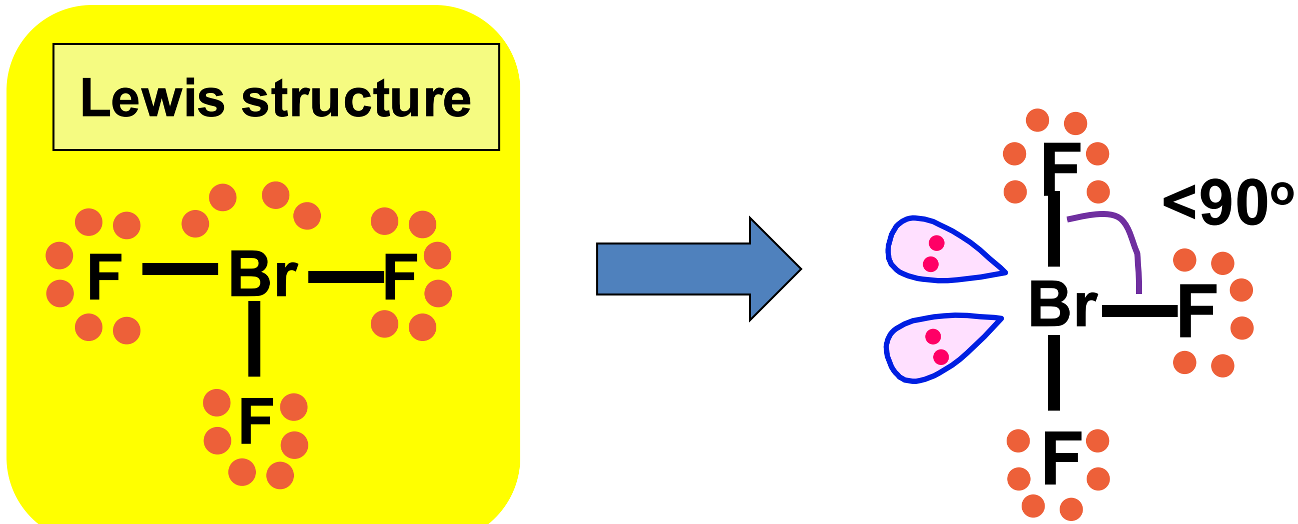
EXAMPLE: SF_4 (sulfur tetrafluoride)





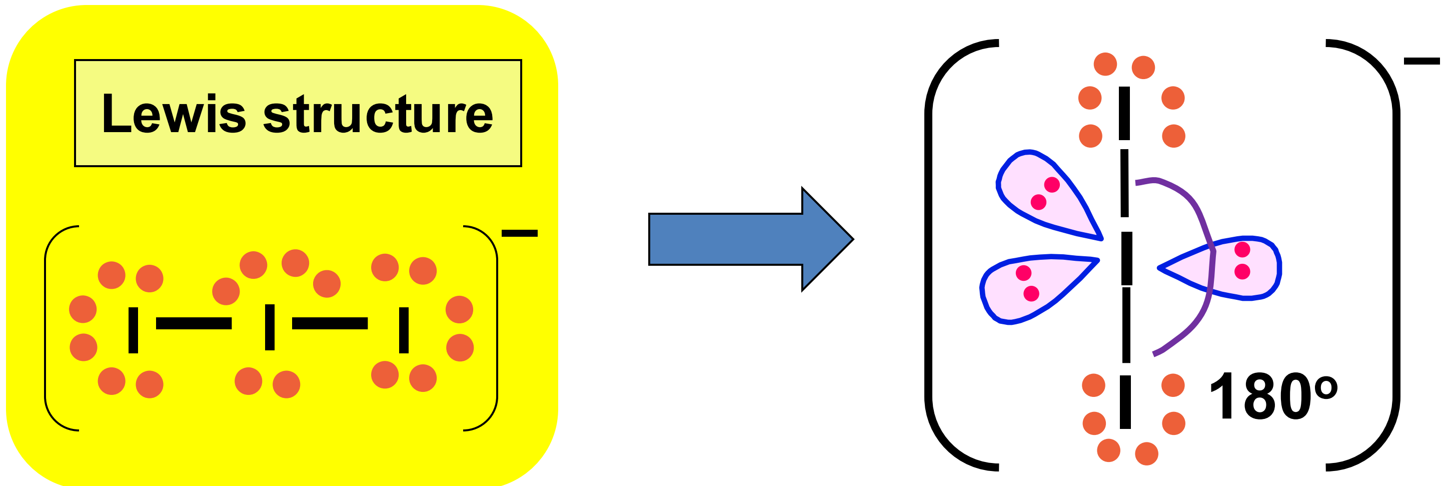
- e⁻ groups arrangement ☞ trigonal bipyramidal
- Molecular shape ☞ see saw (distorted tetrahedral)
  bond angle: < 120°, < 90°

EXAMPLE: BrF_3 (bromine trifluoride)



- e⁻ groups arrangement ☞ trigonal bipyramidal
- Molecular shape ☞ T–shaped
  bond angle: <90°

EXAMPLE: I_3^- (triiodide ion)

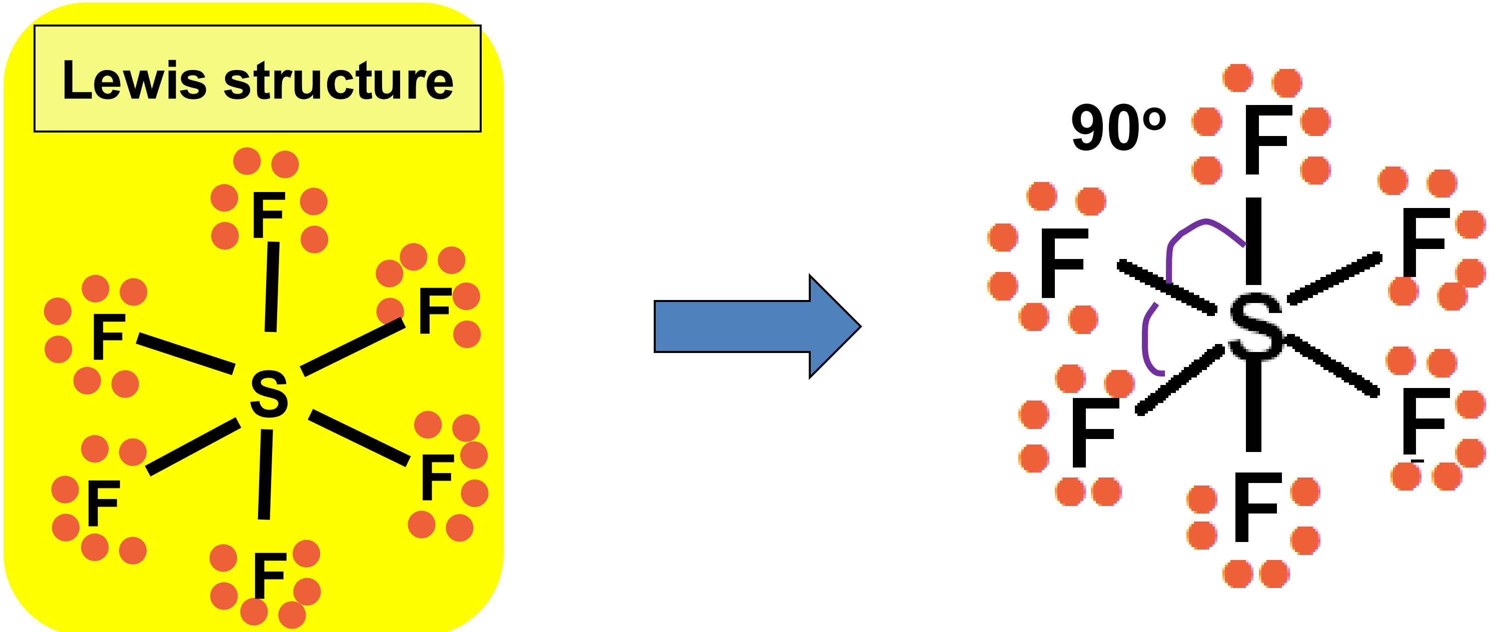


- e⁻ groups arrangement ☞ trigonal bipyramidal
- Molecular shape ☞ linear

  bond angle: 180°

SIX ELECTRON GROUPS

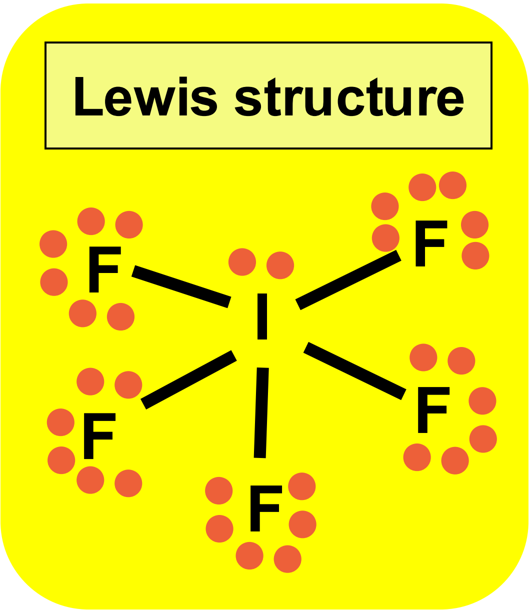
EXAMPLE: SF_6 (sulfur hexafluoride)

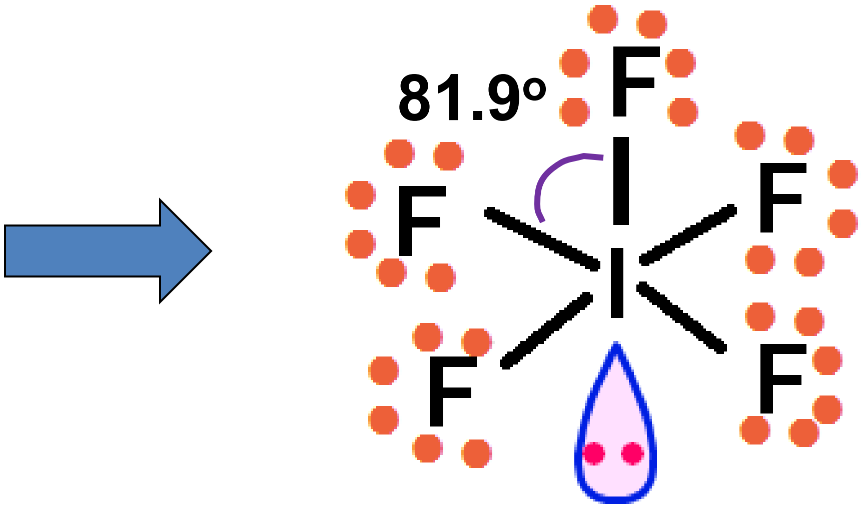


- ❑ e^- groups arrangement ☞ octahedral
- ❑ Molecular shape ☞ octahedral

  bond angle: 90°

EXAMPLE: IF_5 (iodine pentafluoride)

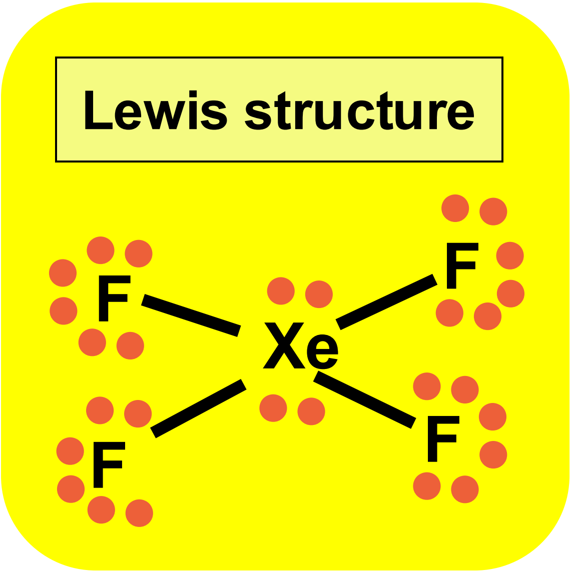


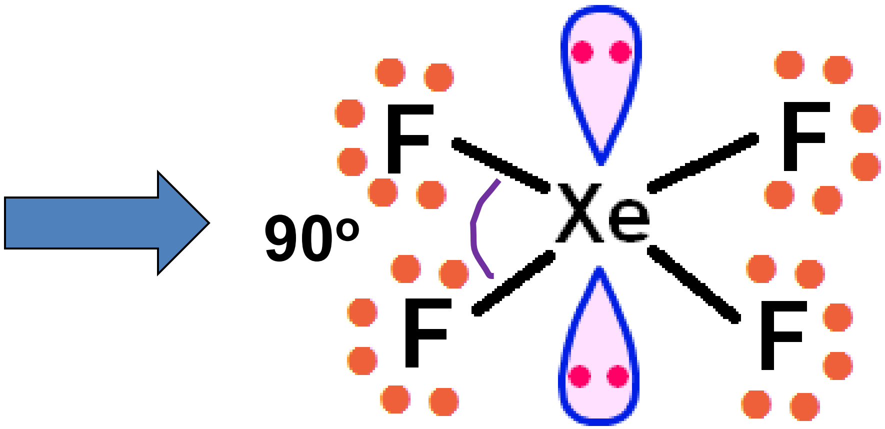


- ❑ e^- groups arrangement ☞ octahedral
- ❑ Molecular shape ☞ square pyramidal

  bond angle: <90°

EXAMPLE: XeF_4 (xenon tetrafluoride)





- ❑ e^- groups arrangement ☞ octahedral
- ❑ Molecular shape ☞ square planar

bond angle: 90º

SUMMARY

e^- group	X	E	e^- group arrangement	molecular shape
2	2	0	linear	linear
3	3	0	trigonal planar	trigonal planar
	2	1	trigonal planar	V–shaped
4	4	0	tetrahedral	tetrahedral
	3	1	tetrahedral	trigonal pyramidal
	2	2	tetrahedral	V–shaped

Note : X - number of bonding pair
E - number of lone pair

e^- g.	X	E	e^- g. arrangement	molecular shape
5	5	0	trigonal bipyramidal	trigonal bipyramidal
	4	1	trigonal bipyramidal	see saw
	3	2	trigonal bipyramidal	T–shaped
	2	3	trigonal bipyramidal	linear
6	6	0	octahedral	octahedral
	5	1	octahedral	square pyramidal
	4	2	octahedral	square planar

Note : X - number of bonding pair
E - number of lone pair

Example 1

4.2

Draw the molecular shape and predict the bond angles (relative to the ideal angles) . What are the electron groups arrangements and the molecular shapes of:

(a) PF_3

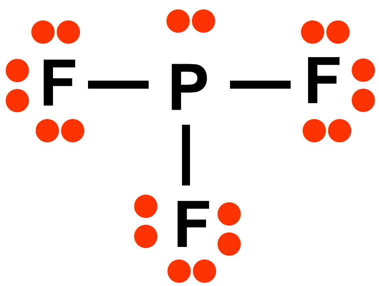
(b) ClF_2^-

Ans: Example 1

4.2

(a) PF_3

Step 1: Draw the Lewis structure:



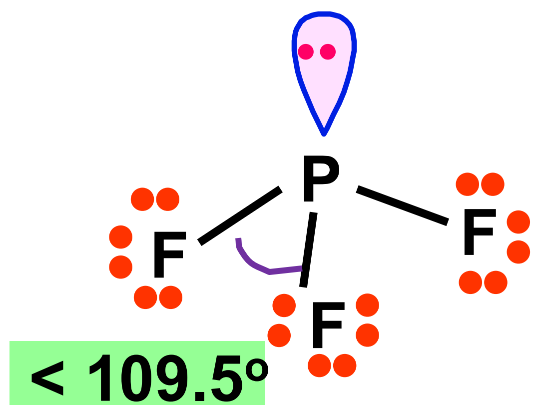
Step 2: Determine the e^- groups arrangement:

- ☞ e^- groups = 4
- ☞ e^- groups arrangement = tetrahedral

Step 3: Predict the bond angle:

- ☞ **without lone pair at P; 109.5°**
- ☞ **According to VSEPR Theory, repulsion of lone pair-bonding pair > bonding pair-bonding pair**
- ☞ **with lone pair at P; < 109.5°**

Step 4: Draw and name the molecular shape:

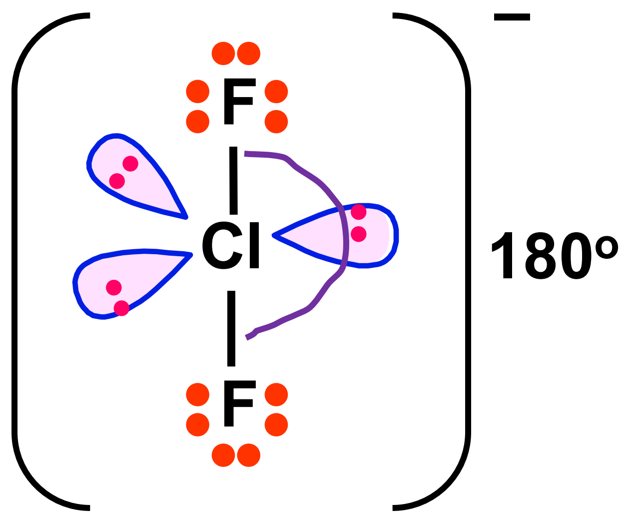


Molecular shape = Trigonal pyramidal

Ans: Example 1

4.2

(b) ClF_2^-



e^- groups arrangement = trigonal bipyramidal

Molecular shape = linear

BOND POLARITY

- Atoms with different electronegativities form polar bonds

- Depicted as a polar arrow:

EXAMPLE:

$\delta+$ $\delta-$
C—Cl

$\delta+$ $\delta-$
C—F

⇩

$\mapsto$
C—Cl

☞ polar bond

$\longmapsto$
C—F

☞ polar bond

C—C

☞ nonpolar bond

Br—Br

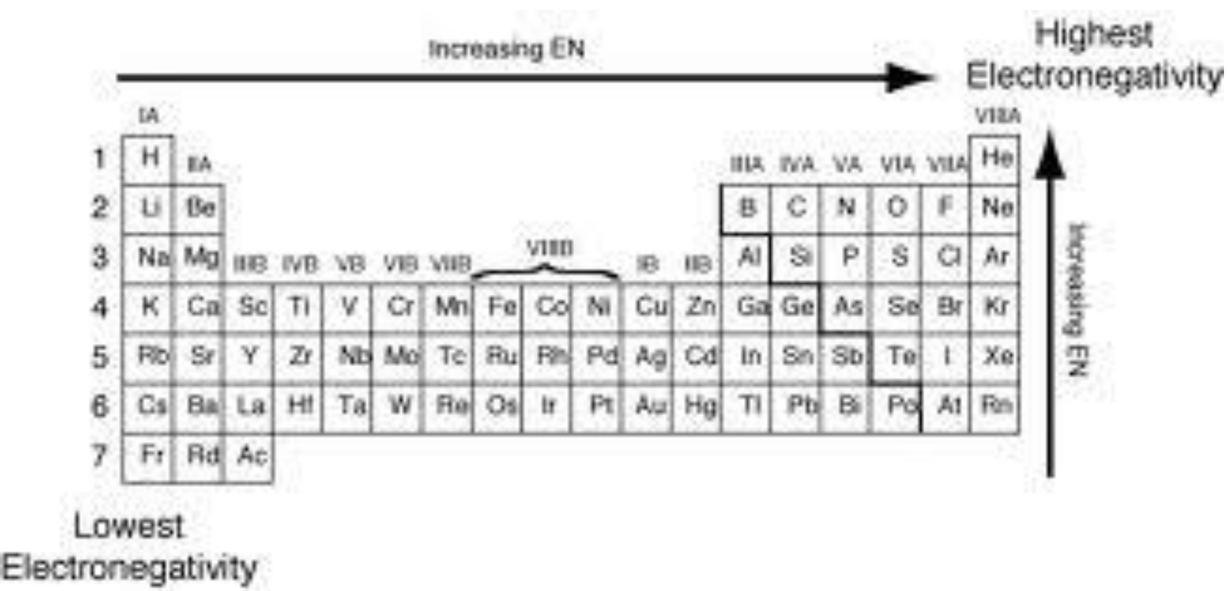
☞ nonpolar bond

Example 2

4.2

(a) Use a polar arrow to indicate the polarity of each bond: N–H, F–N, I–Cl

(b) Rank the following bonds in order of increasing polarity: H–N, H–O, H–C



Ans: Example 2

4.2

(a)

N—H

N is more electronegative than H

F—N

F is more electronegative than N

I—Cl

Cl is more electronegative than I

Ans: Example 2

4.2

(b) **Bond polarity:**

H–N , H–O , H–C

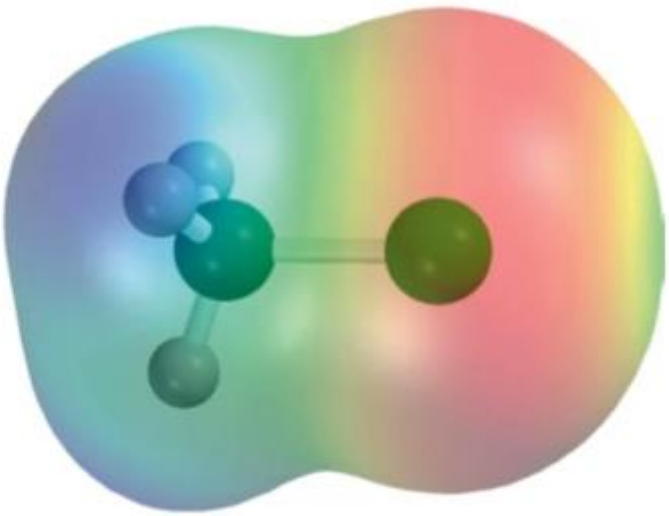
So, Bond polarity:

H–C < H–N < H–O

MOLECULAR POLARITY

- Net imbalanced of charge
- e$^-$ rich regions (δ–) and e$^-$ poor regions (δ+)

electron poor region



electron rich region

Chloromethane (CH_3Cl)

DIPOLE MOMENT (μ)

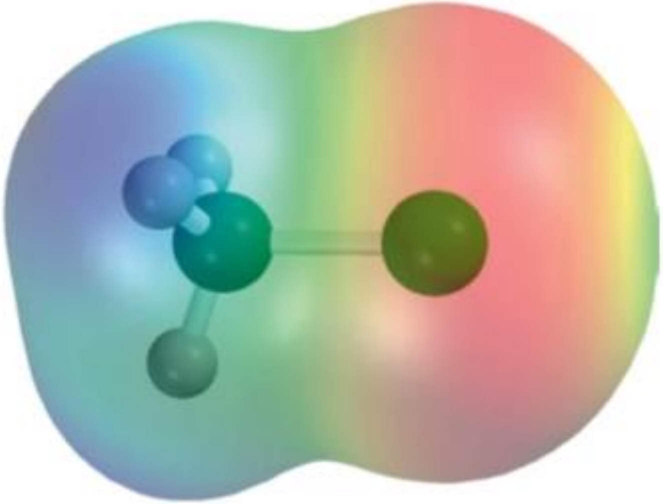
- **Quantitative measure of molecular polarity**

$\mu = Q \times r$

Q: charge

r: distance between charges

1 D (Debye) = 3.36×10^{-30} C m



Chloromethane (CH_3Cl)

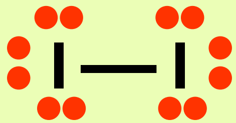
Peter Joseph William Debye (1884-1966). American chemist and physicist.

RESULTANT DIPOLE MOMENT

- Determined by molecular shape and bond polarity

- $\mu \neq 0$ ☞ polar
- $\mu = 0$ ☞ nonpolar

EXAMPLE:

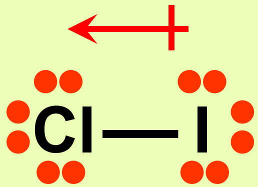


Δelectronegativity = 0

$\mu = 0$

I_2 is a nonpolar molecule

EXAMPLE:



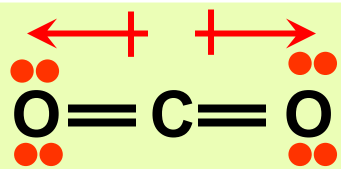
$\mu \neq 0$

Δelectronegativity $\neq$ 0

$\mu \neq 0$

ICl is a polar molecule

EXAMPLE:



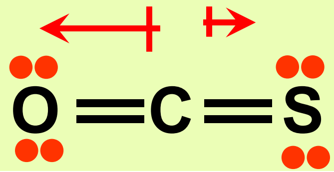
CO_2 shape: linear

The C–O bonds are polar but the two bond dipoles cancel each other

Resultant dipole moment, $\mu = 0$

CO_2 is a nonpolar molecule

EXAMPLE:



resultant $\mu \neq 0$

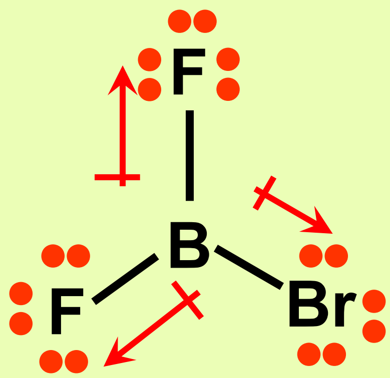
OCS shape: linear

The C–O and C–S bond are polar and the two bond dipoles do not cancel each other

Resultant dipole moment, $\mu \neq 0$

OCS is a polar molecule

EXAMPLE:



resultant $\mu \neq 0$

BF_2Br shape: trigonal planar

The B–F and B–Br bond are polar and the three bond dipoles do not cancel each other

Resultant dipole moment, $\mu \neq 0$

BF_2Br is a polar molecule

Example 3

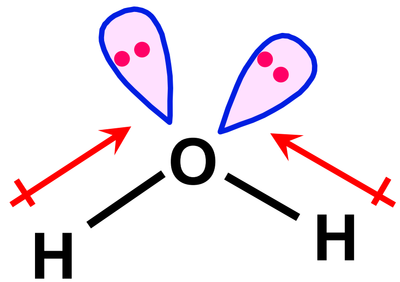
4.2

Which of the following molecules have a dipole moment (polar molecule) ? Explain.

H_2O or CCl_4

Ans: Example 3

4.2



resultant $\mu \neq 0$

Shape = bent

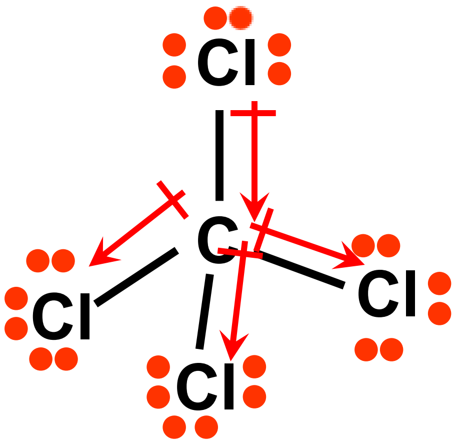
The O–H bonds are polar and the two bond dipoles do not cancel each other

Resultant dipole moment, $\mu \neq 0$

H_2O is a polar molecule

Ans: Example 3

4.2



resultant μ = 0

Shape = linear

The C–Cl bonds are polar but all four bond dipoles cancel each other

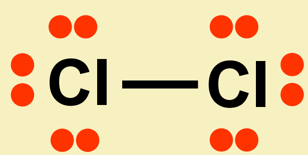
Resultant dipole moment, μ = 0

CCl_4 is a nonpolar molecule

SUMMARY

- **A molecule will be nonpolar if:**
  - **The bonds are nonpolar**

EXAMPLE:

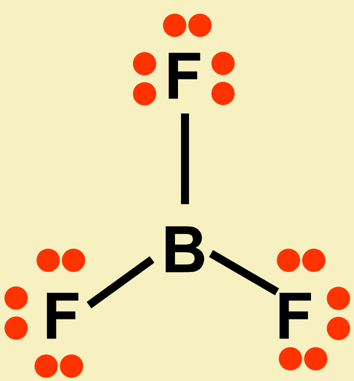


a nonpolar molecule

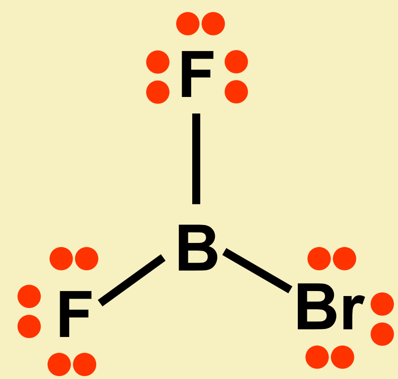
- A molecule will be nonpolar if:

    - No lone pair in the central atom and all the surrounding atoms are the same (the molecular shape is a basic shape)

EXAMPLE:



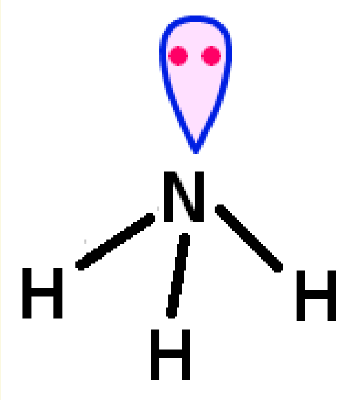
a nonpolar molecule

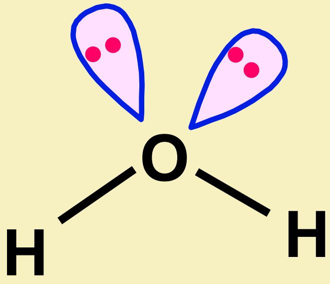


a polar molecule

- A molecule in which the central atom has lone pair e^- will usually be polar with few exceptions

EXAMPLE:

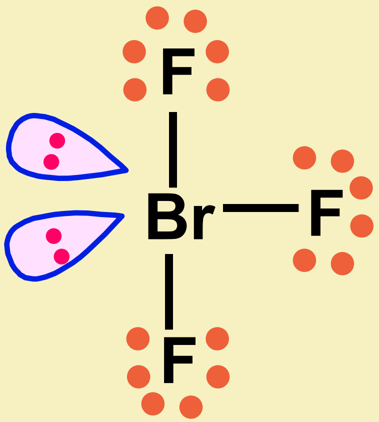




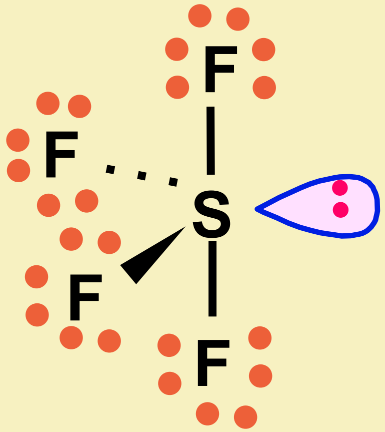
a polar molecule

a polar molecule

EXAMPLE:

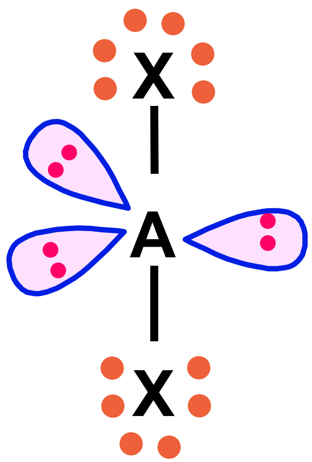


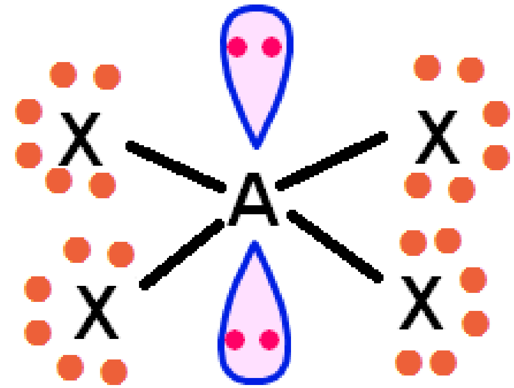
a polar molecule



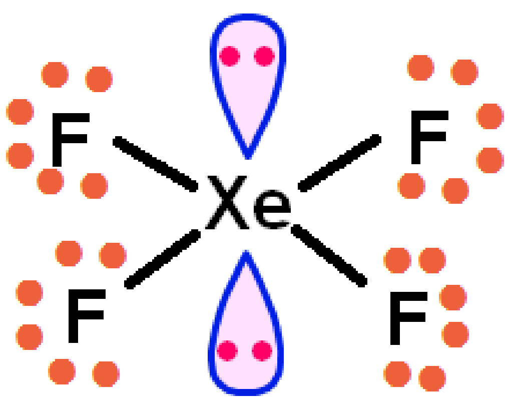
a polar molecule

- **Exception:**





EXAMPLE:



a non polar

Example 4

4.2

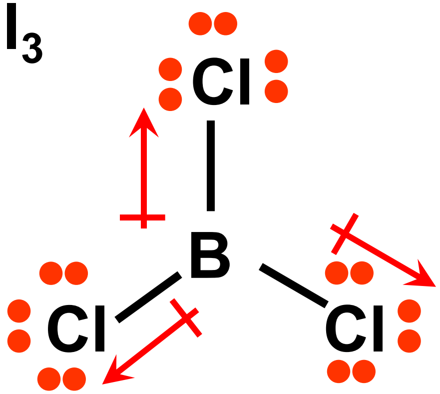
Predict whether each of the following molecules is polar and show the direction of bond polarity and net dipole moment.

(a) Boron trichloride, BCl_3
(b) Hydrogen bromide, HBr

Ans: Example 4

4.2

(a) BCl_3

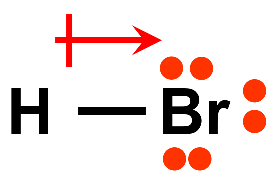


B–Cl bonds are polar but the two bond dipoles cancel each other

$\mu = 0$

BCl_3 is a nonpolar molecule

(b) HBr



H–Br bond is polar and the bond dipoles do not cancel each other

$\mu \neq 0$

HBr is a polar molecule

4.3 ORBITAL OVERLAP AND HYBRIDIZATION

Teaching and learning outcomes

At the end of the lesson, student should be able to

4.1	**Orbital overlap and hybridization**
a)	Illustrate the formation of sigma (σ) and pi (π) bonds from overlapping of orbitals. (C4)
b)	Describe the formation of hybrid orbitals of a central atom: sp^3, sp^2, sp, sp^3d, sp^3d^2. (C1, C2)
c)	Illustrate the hybridisation of the central atom and the overlapping of orbitals in molecules. (C4)

CHAPTER 4.3 : OVERVIEW

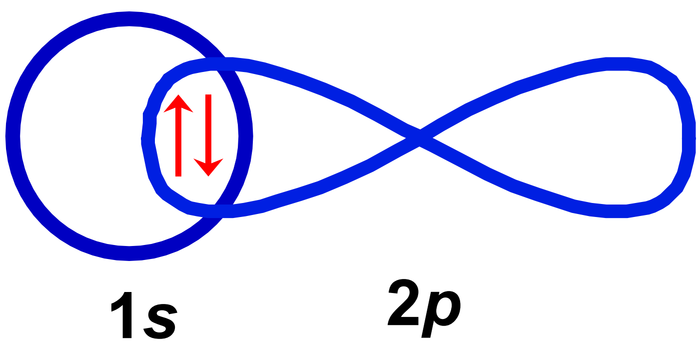
Orbital Overlap & Hybridization

- Describe the formation of sigma (σ) and pi (π) bonds from overlapping orbitals
- Draw and explain the formation of hybrid orbitals for central atom
  - *sp, sp^2, sp^3, sp^3d & sp^3d^2*
- Draw orbitals overlap and label sigma (σ) and pi (π) bonds of a molecule

VALENCE BOND (VB) THEORY

- According to this theory, **a covalent bonds** are between two atoms is **formed** when **a pair of electrons is shared** by two **overlapping atomic orbitals**

EXAMPLE:



DIRECT ORBITALS OVERLAP

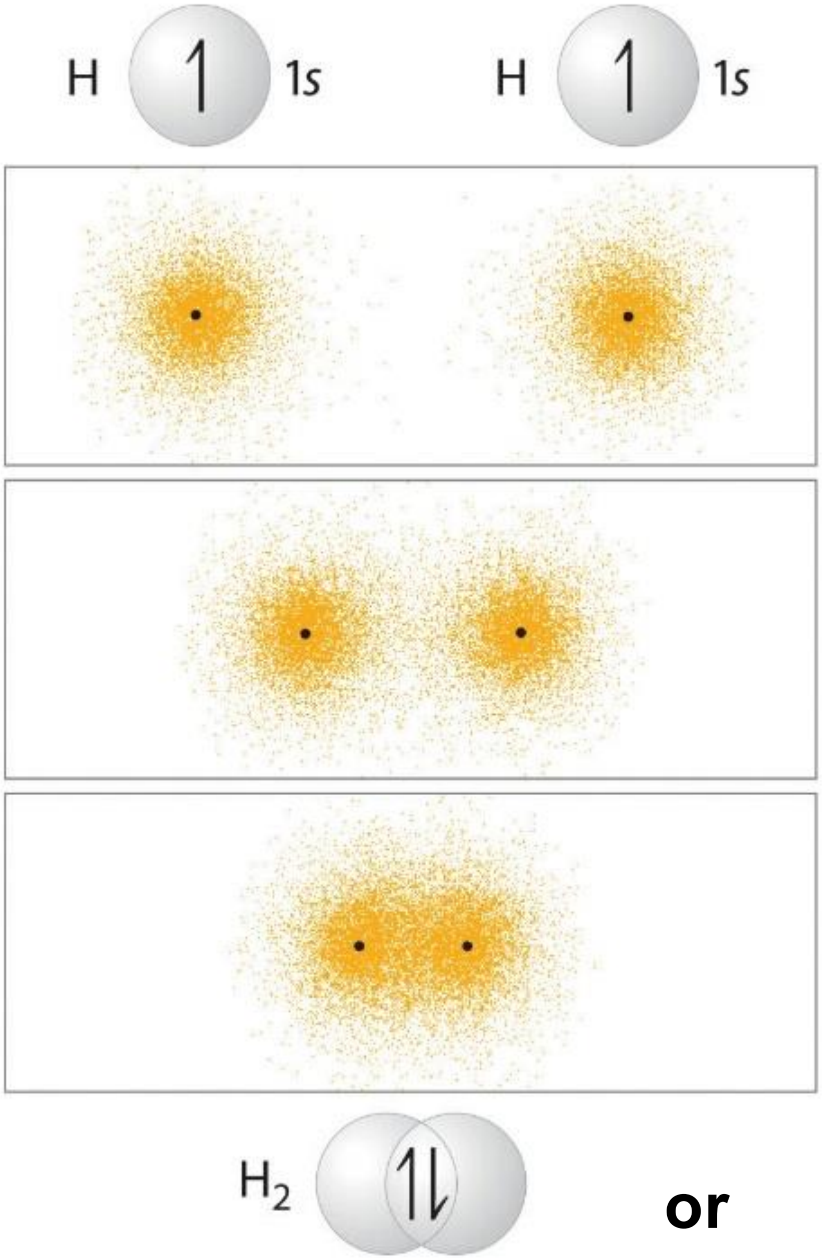
- Atoms in simple molecules or ions such as H_2, HF, N_2, normally use pure *s* and/or *p* orbitals in forming covalent bonds

EXAMPLE:

s orbital overlaps with *s* orbital

p orbital overlaps with *p* orbital

s orbital overlaps with *p* orbital



Each H atom has one unpaired electron in its valence shell

Both H atoms are far apart

Atoms begin to interact as they move closer to each other

As the two atoms approach each other, their 1s orbitals overlap and the electron pair spread out over both orbital

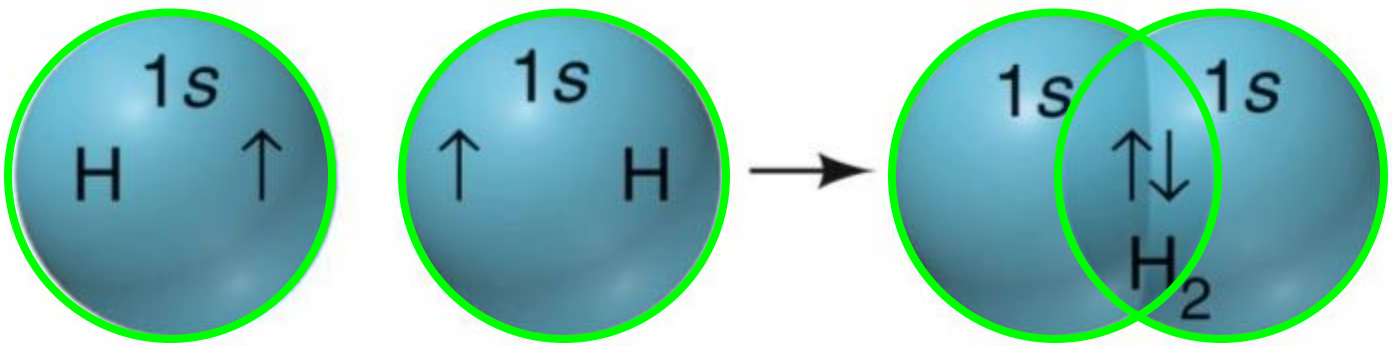
or

H—H → Covalent bond

EXAMPLE: H_2 (hydrogen molecule)

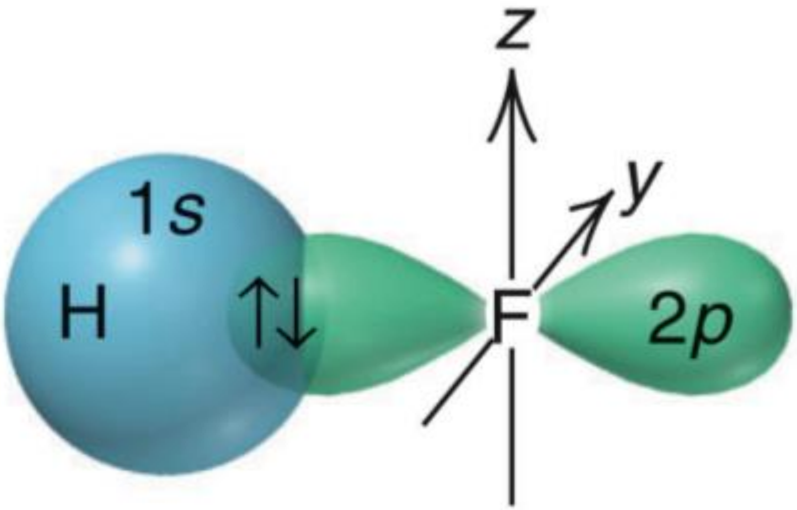
H $1s^1$ $\boxed{\uparrow}$ $1s$

H $1s^1$ $\boxed{\uparrow}$ $1s$



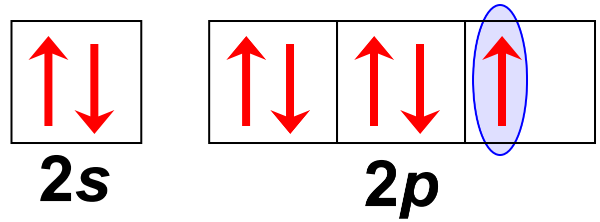
EXAMPLE: HF (hydrogen fluoride)

H	$1s^1$	↑				
		1*s*				
F	$1s^2\ 2s^2\ 2p^5$		↑↓	↑↓	↑↓	↑
			2*s*	2*p*		

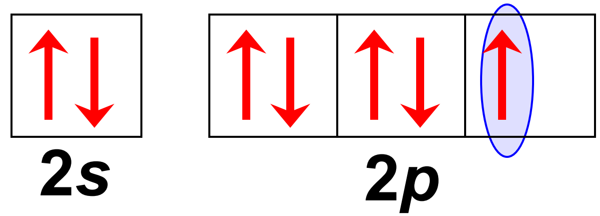


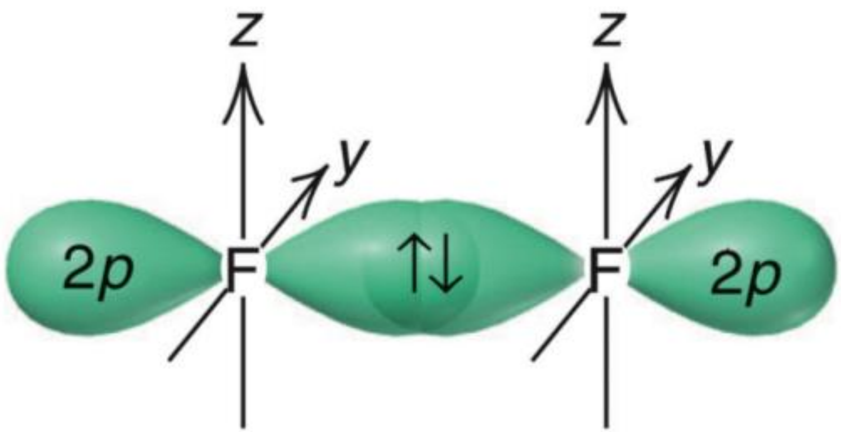
EXAMPLE: F_2 (fluorine molecule)

F $1s^2\ 2s^2\ 2p^5$



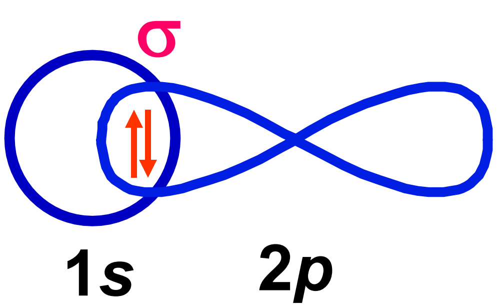
F $1s^2\ 2s^2\ 2p^5$



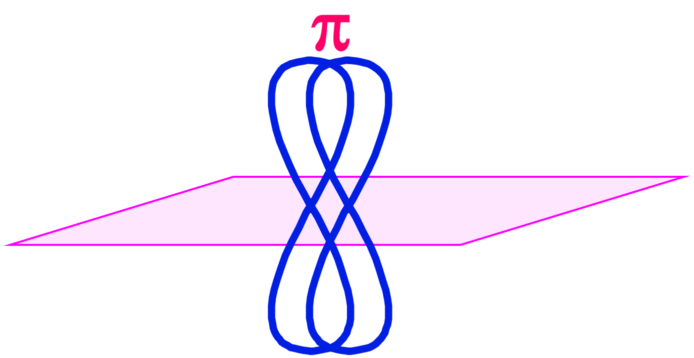


MODE OF OVERLAP

- Two types:
  - End–to–end ☞ sigma (σ) bond
  - Side–to–side ☞ pi (π) bond



end–to–end overlap



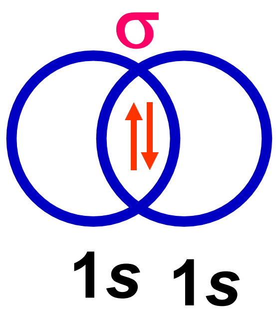
side–to–side overlap
(sideway)

SIGMA (σ) BOND

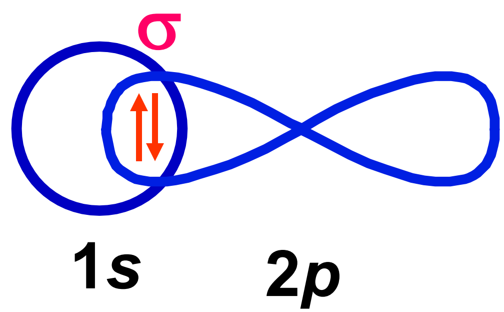
- Resulting from end–to–end overlap
- Has highest e^- density along the bond axis
- Allow free rotation
- All single bonds are σ bond

EXAMPLE:

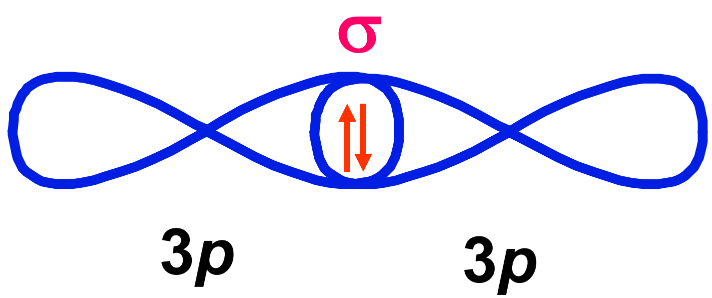
Overlaping between original orbital



H$\overset{\sigma}{—}$H bond



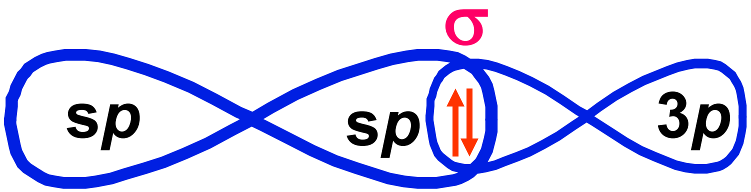
H$\overset{\sigma}{—}$F bond

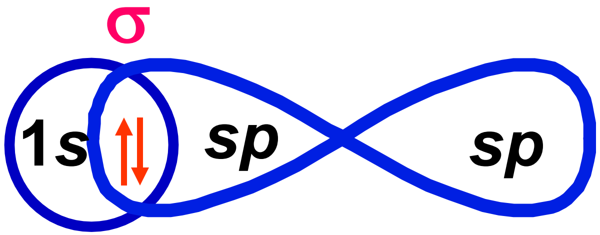


Cl$\overset{\sigma}{—}$Cl bond

EXAMPLE:

Overlaping with hybrid orbital

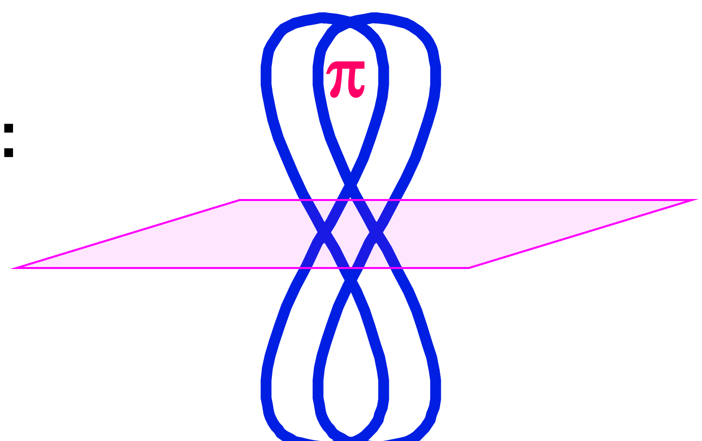




PI (π) BOND

- Resulting from **side–to–side** overlap (sideway)
- Has **two regions** of e^- density
  - ☞ one above and one below the **σ–bond** axis

EXAMPLE:

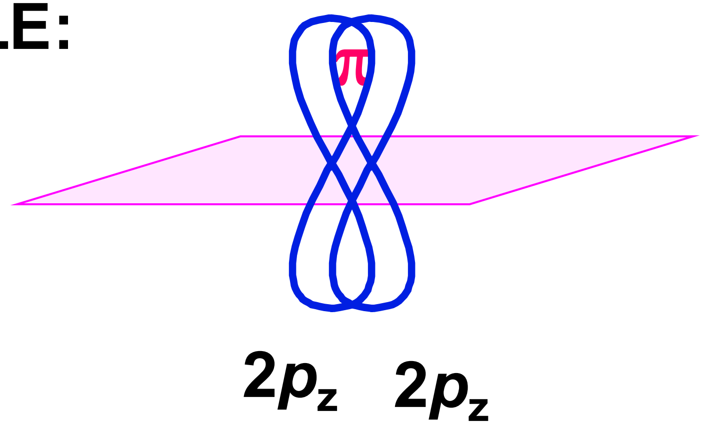


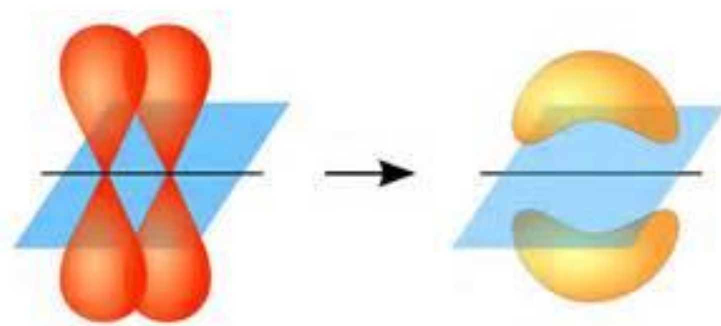
$2p_z$

- One π **bond** hold **two e$^-$** that move through both regions of the bond

- π bond **restricts the rotation**

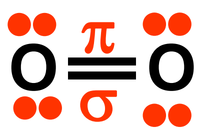
EXAMPLE:





- **Double bond consists of one σ bond and one π bond**

EXAMPLE:

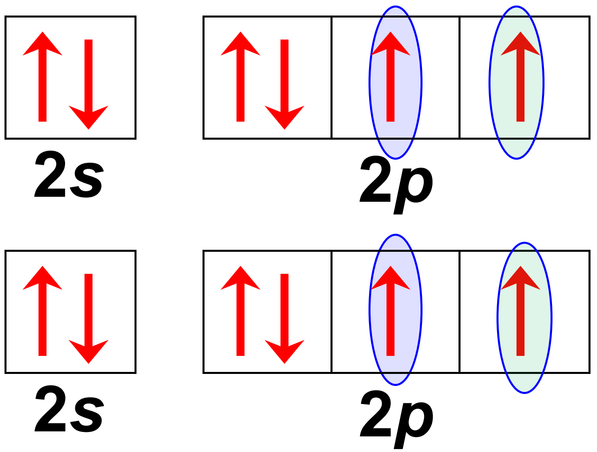


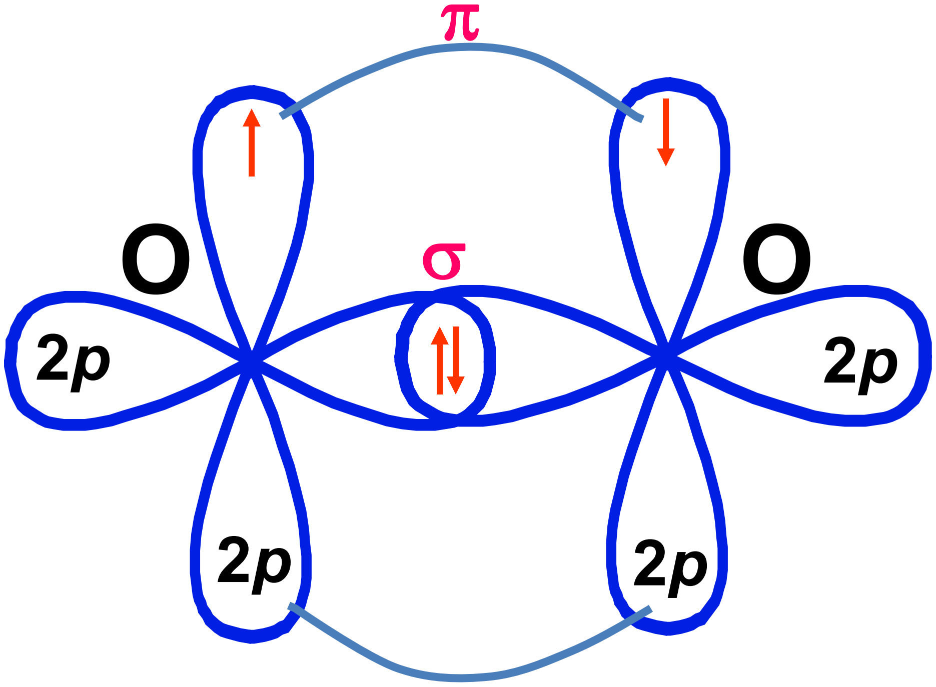
☞ **O_2 has one π bond and one σ bond**

EXAMPLE: O_2 (oxygen molecule)

O $1s^2\ 2s^2\ 2p^4$

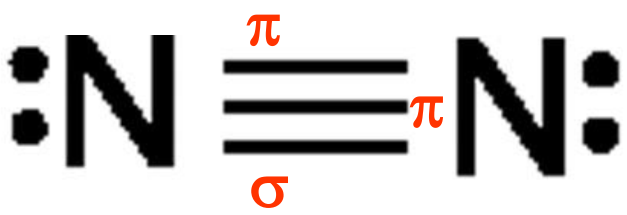
O $1s^2\ 2s^2\ 2p^4$





- **Triple bond** **consists of** **one σ** **bond and two π** **bond**

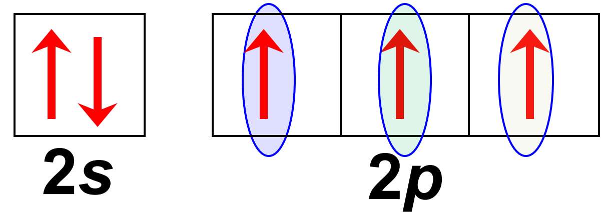
EXAMPLE:



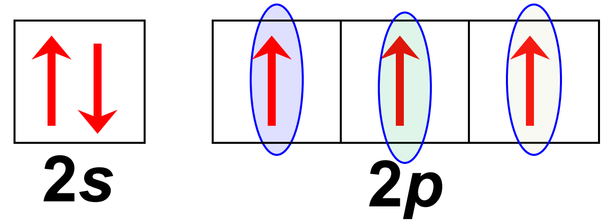
☞ **N_2 has two π bonds and one σ bond**

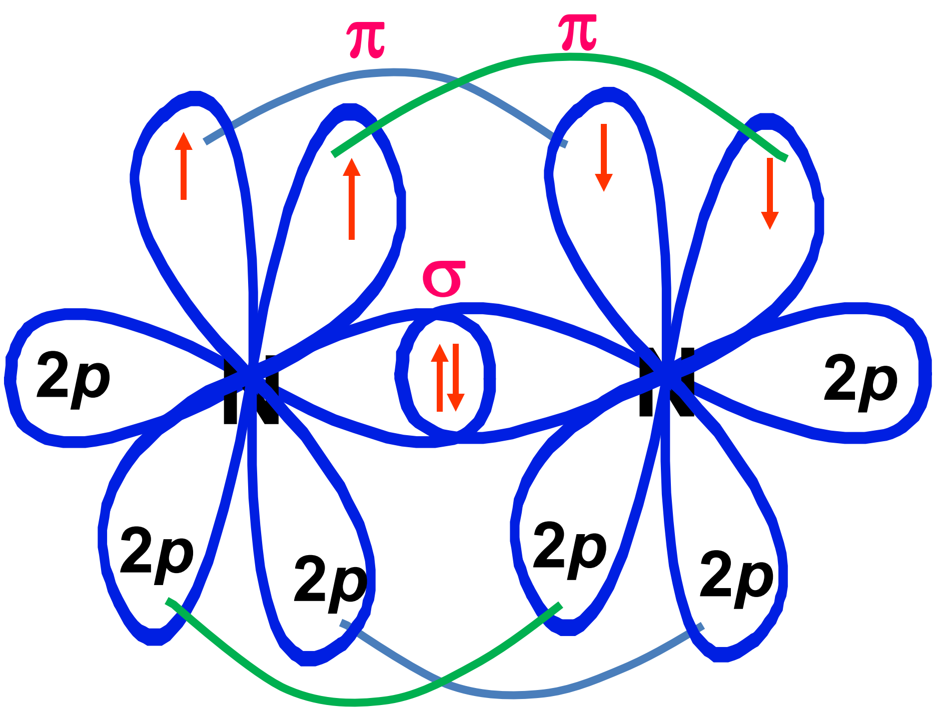
EXAMPLE: N_2 (nitrogen molecule)

N $1s^2\ 2s^2\ 2p^4$



N $1s^2\ 2s^2\ 2p^4$

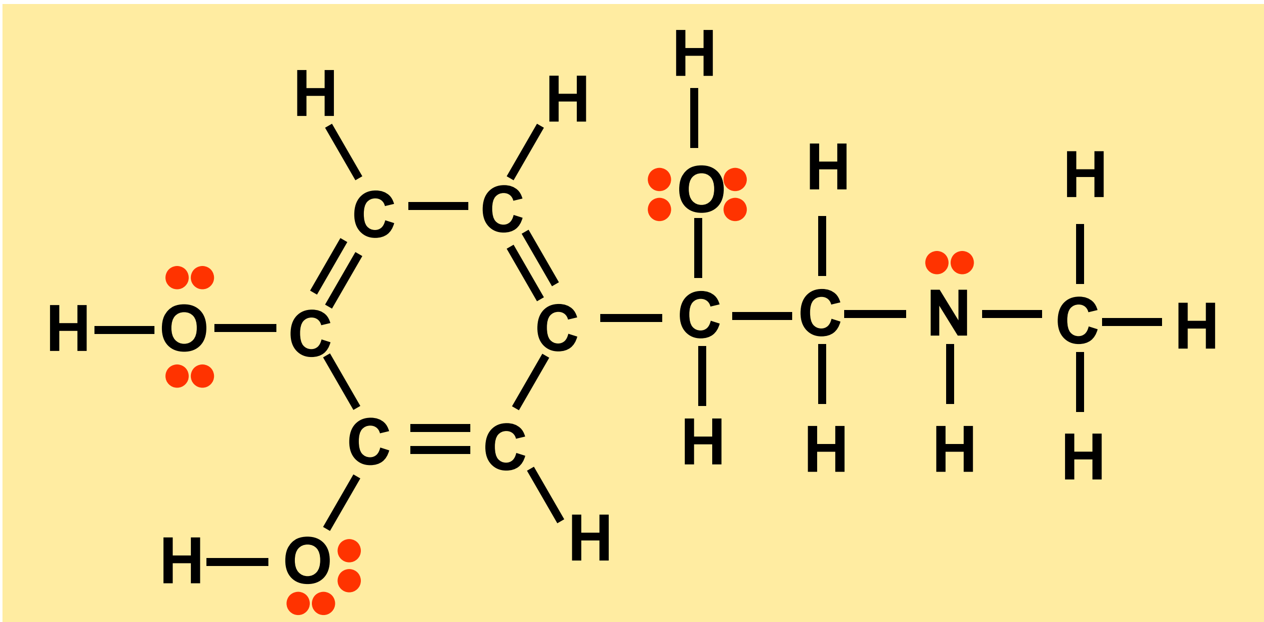




Example 1

4.3

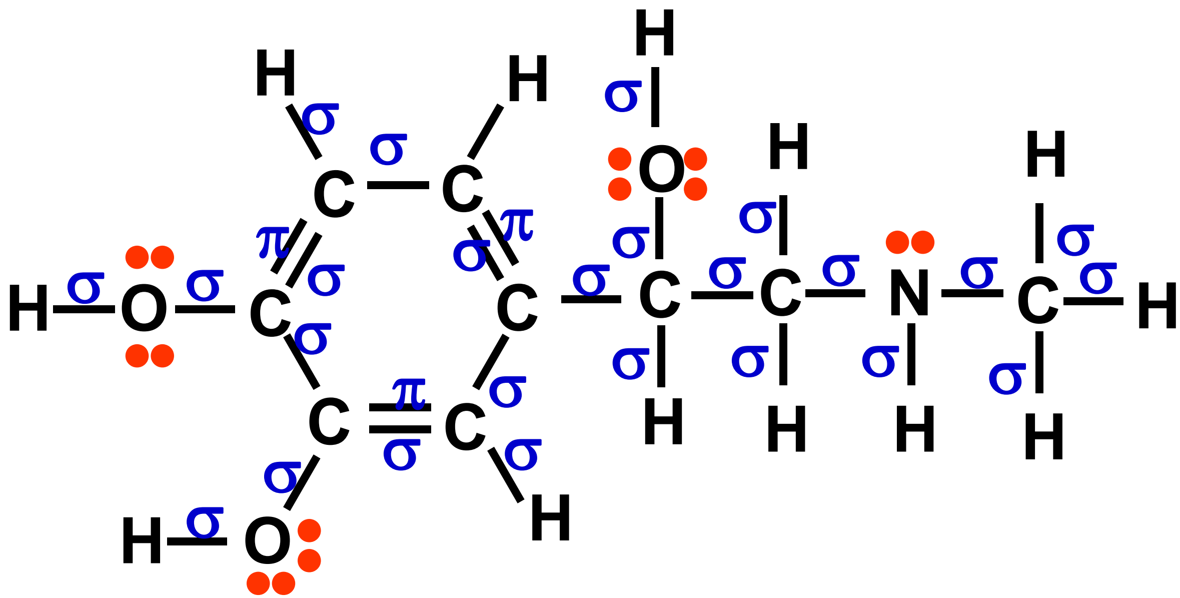
Adrenaline has the following Lewis structure:



How many σ and π bonds are in the molecule ?

Ans: Example 1

4.3



Total: **26 σ bonds and 3 π bonds**

HYBRIDIZATION

- **Mixing of two or more atomic orbitals to form a new set of equivalent hybrid orbitals**

- **The spatial orientation of the new orbitals is caused more stable bonds and are consistent with the observed molecular shape**

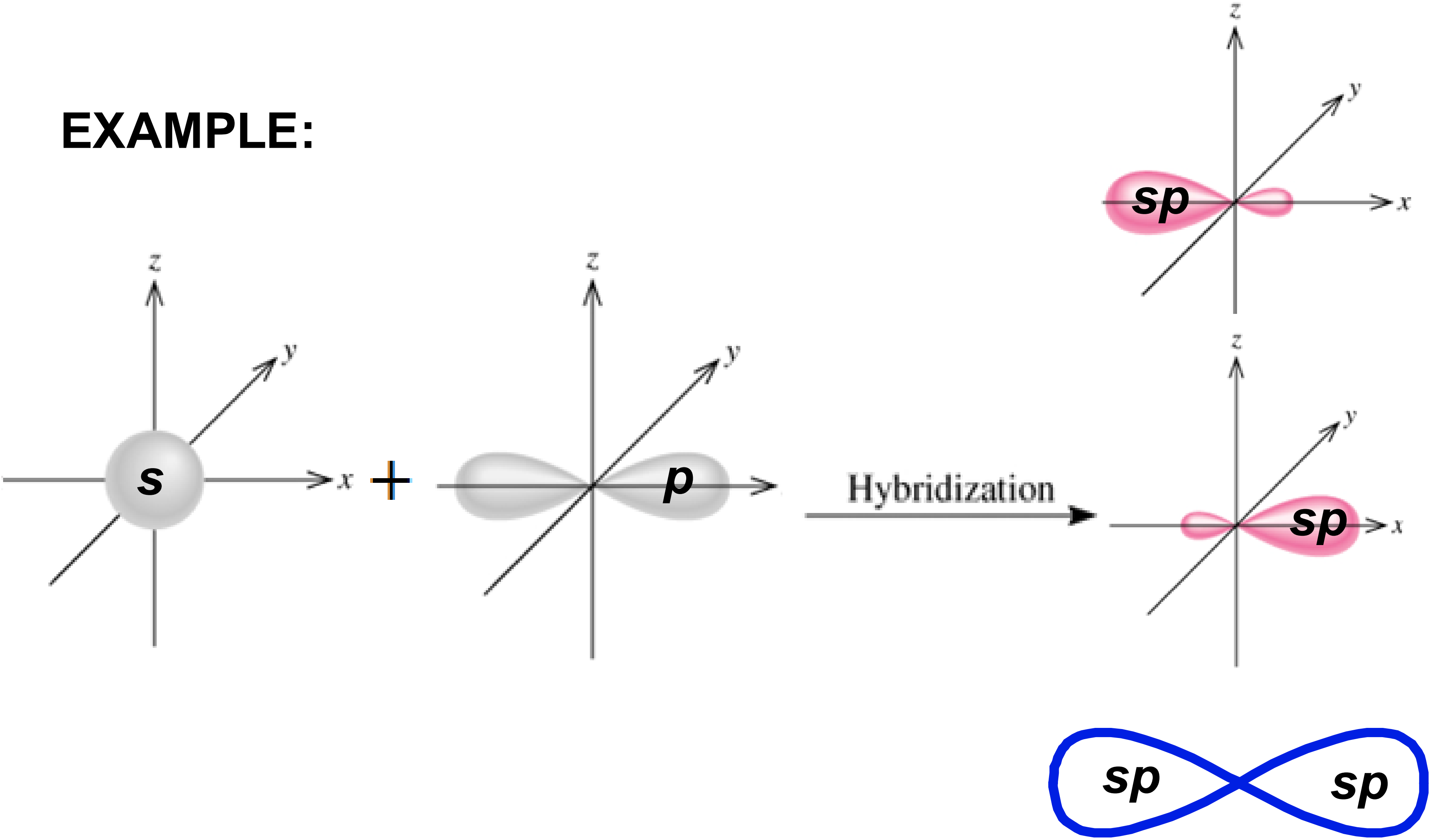
- **Number** of hybrid orbitals obtained **equals** the number of **atomic orbitals mixed**

- **Type** of hybrid orbitals obtained varies with the types of **atomic orbitals mixed**

TYPES OF HYBRID ORBITALS

Hybrization	Type
s + p	sp
s + p + p	sp^2
s + p + p + p	sp^3
s + p + p + p + d	sp^3d
s + p + p + p + d + d	sp^3d^2

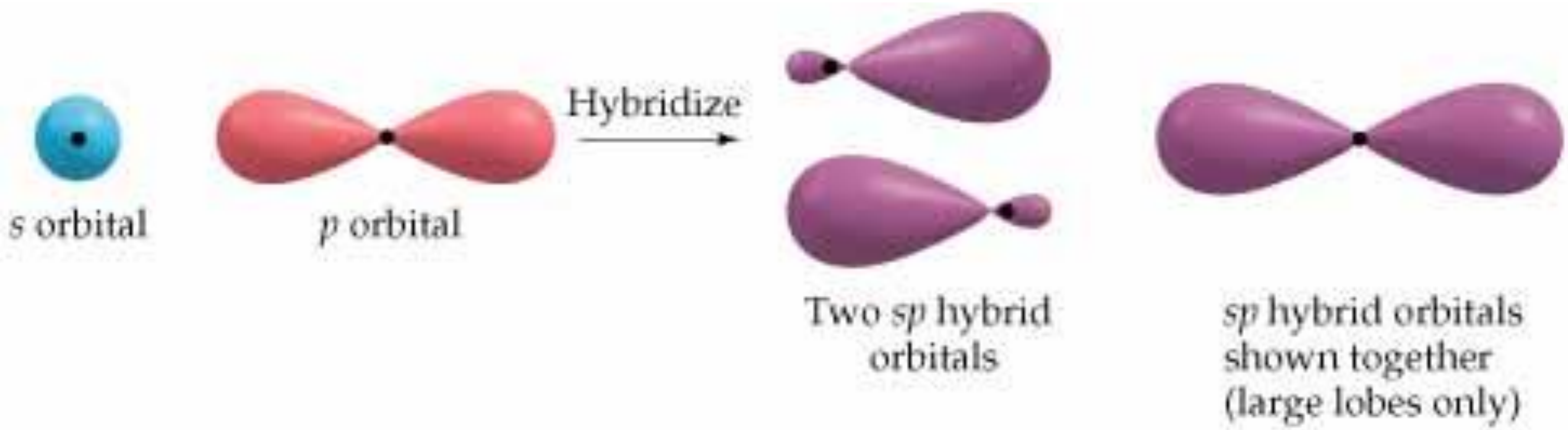
EXAMPLE:

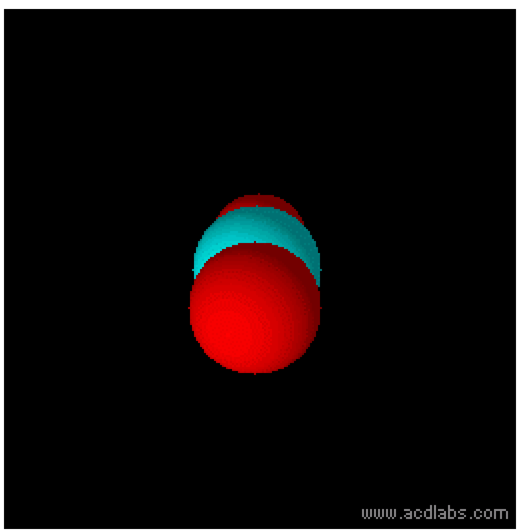


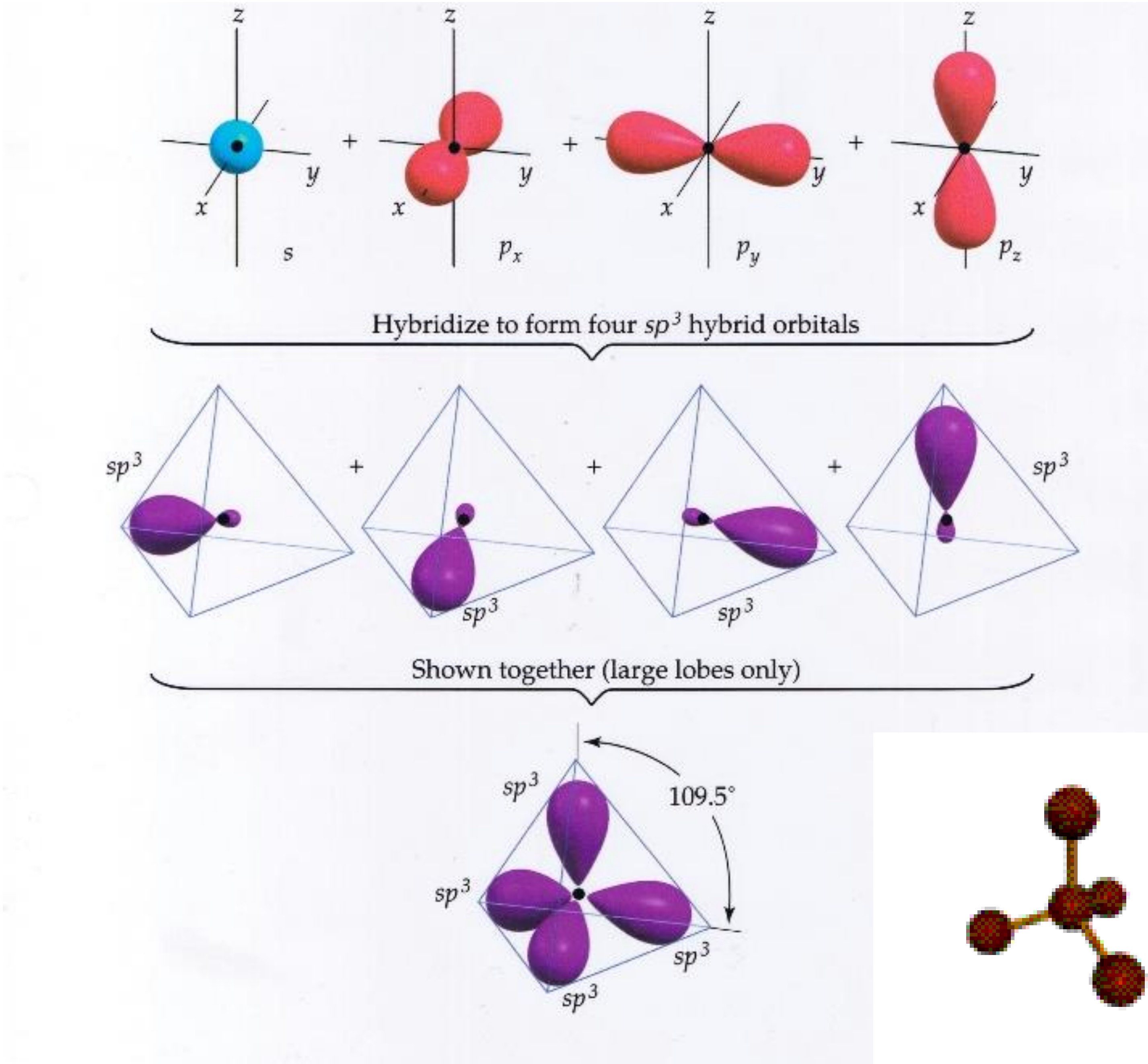
s orbital

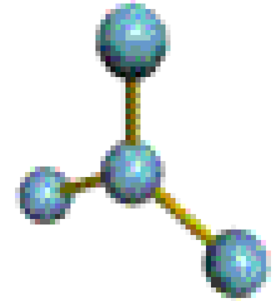
p orbital

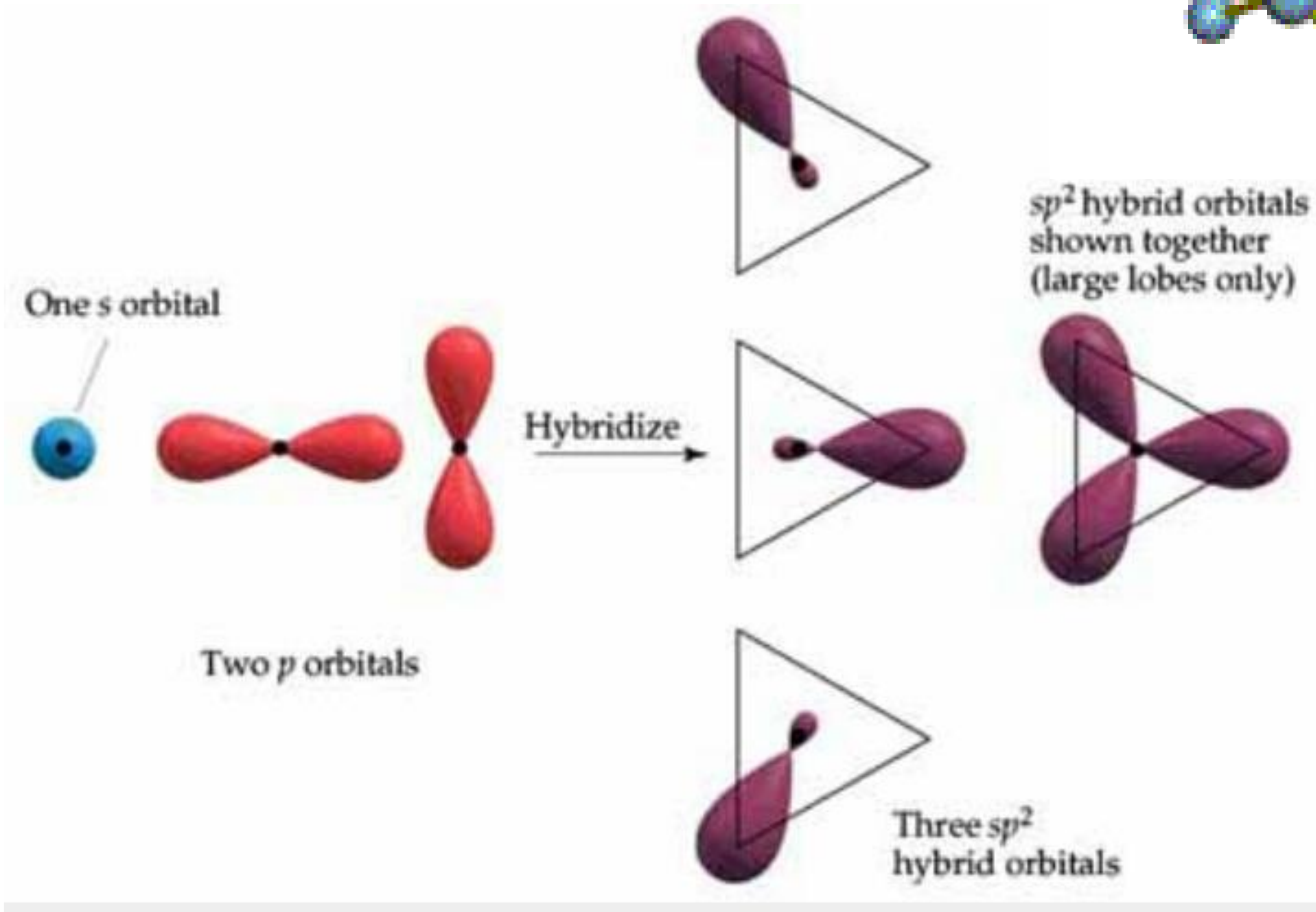
sp orbitals

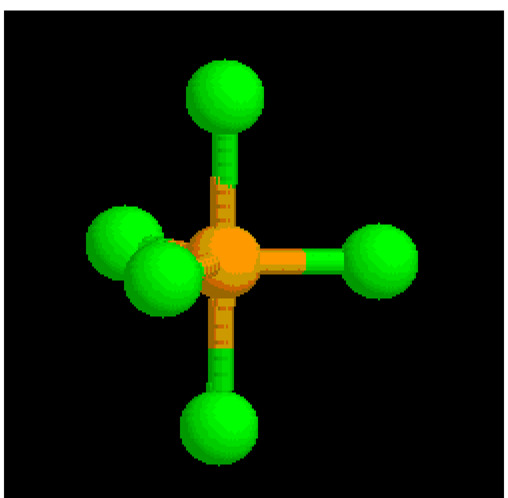

Hybridize
s orbital
p orbital
Two sp hybrid orbitals
sp hybrid orbitals shown together (large lobes only)

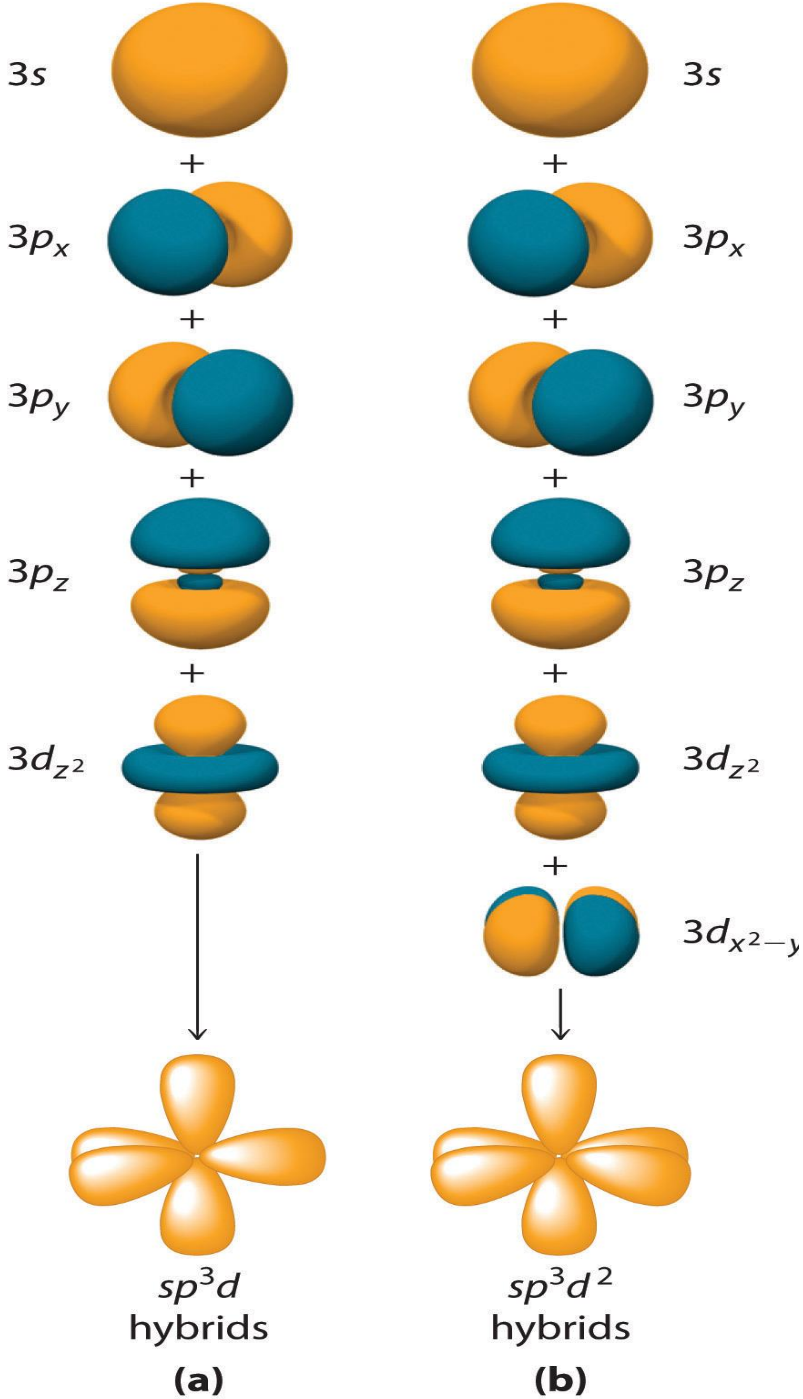

www.acdlabs.com

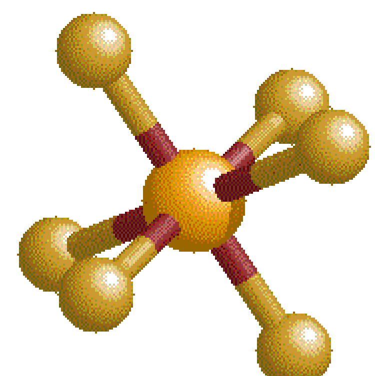

z
x
y
s
+
p_x
p_y
p_z
Hybridize to form four sp^3 hybrid orbitals
sp^3
Shown together (large lobes only)
109.5°




One s orbital
Two p orbitals
Hybridize
sp^2 hybrid orbitals shown together (large lobes only)
Three sp^2 hybrid orbitals




3s
3p_x
3p_y
3p_z
3d_z2
3d_x2−y
sp3d
hybrids
sp3d2
hybrids
(a)
(b)

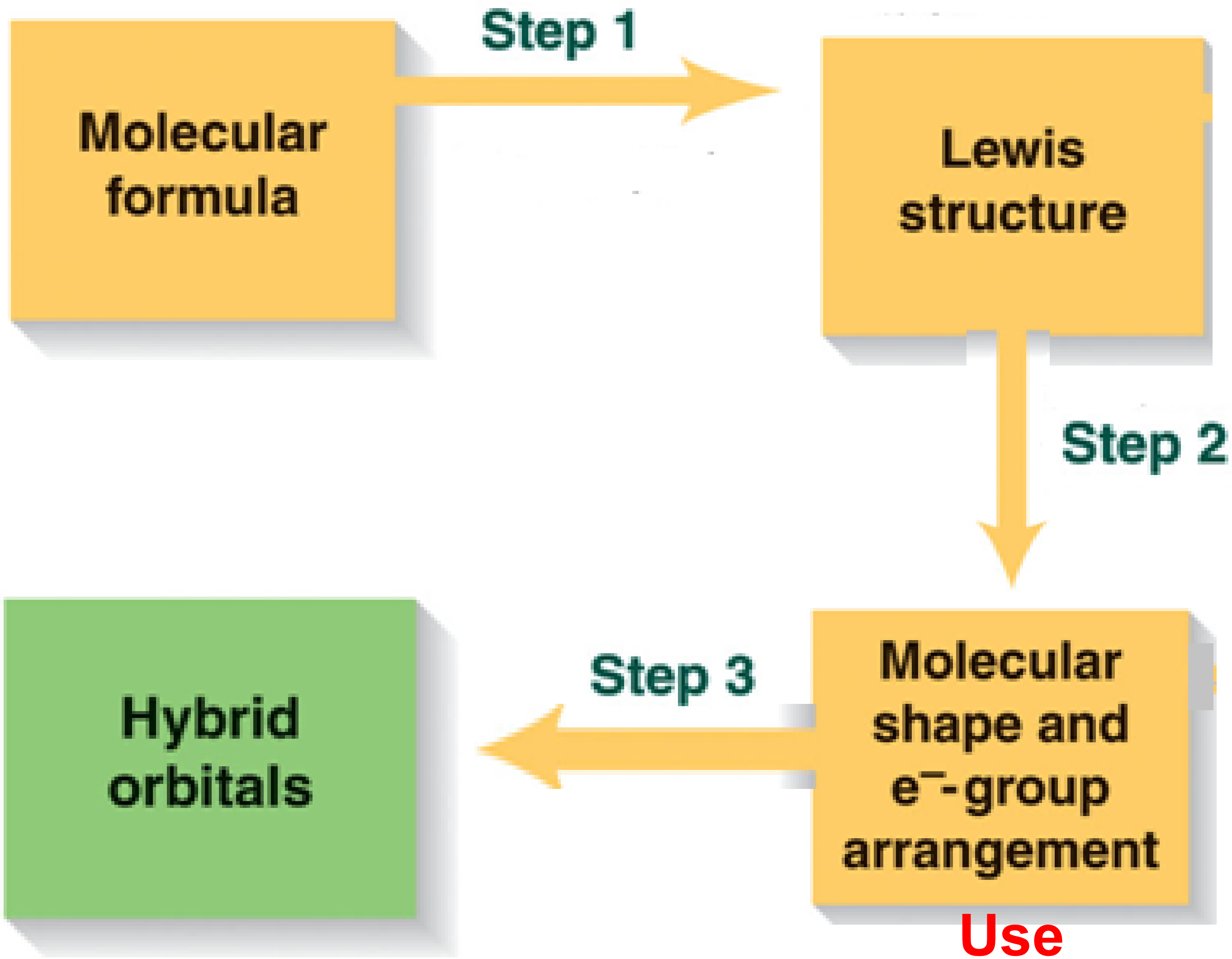


DETERMINING TYPE OF HYBRID ORBITALS

1. Draw **Lewis structure**

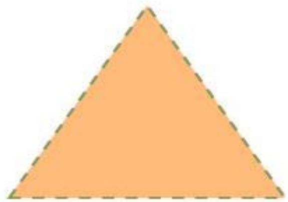
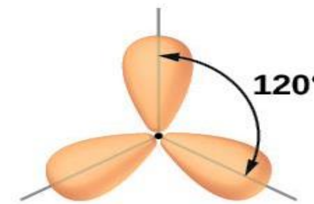
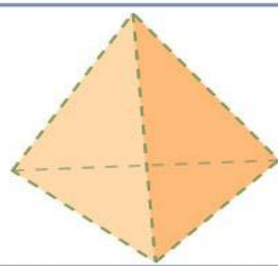
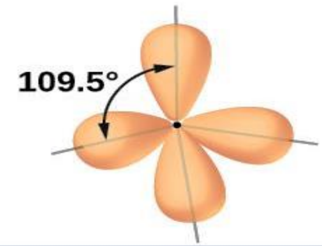
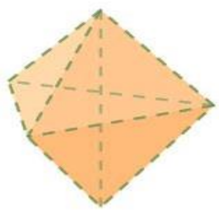
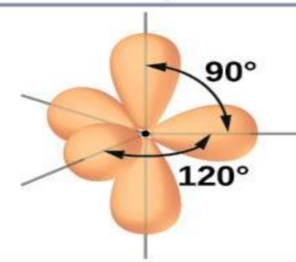
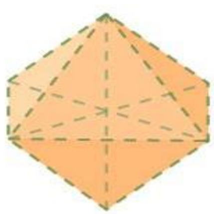
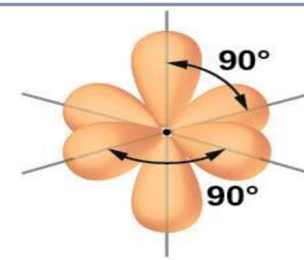
2. Predict the **e^- groups** arrangement using **VSEPR** model

3. Deduce the **hybridization** of the central atom by matching the **arrangement** of the e^- groups with the hybrid orbitals



Use VSEPR theory

e^- Group	e^- Group Arrangement	Type of hybridization
2	linear	*sp*
3	trigonal planar	sp^2
4	tetrahedral	sp^3
5	trigonal bipyramidal	sp^3d
6	octahedral	sp^3d^2

e^- group	e^- group arrangement		Hybrid orbitals	
2		linear	sp	
3		trigonal planar	sp^2	
4		tetrahedral	sp^3	
5		trigonal bipyramidal	sp^3d	
6		octahedral	sp^3d^2	

Example 2

4.3

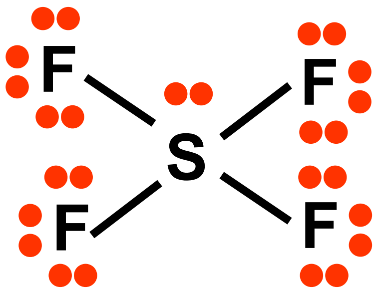
Determine the hybridization state of the central (underlined) atom in each of the following molecules:

(a) SF_4
(b) BeF_2
(d) $SiCl_4$
(e) XeF_4

Ans: Example 2

4.3

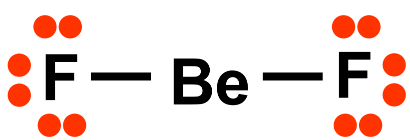
(a) SF_4 Lewis structure:



S: Electron–groups = 5

Hybridization of S = ***sp*^{3}*d***

(b) BeF_2 Lewis structure:



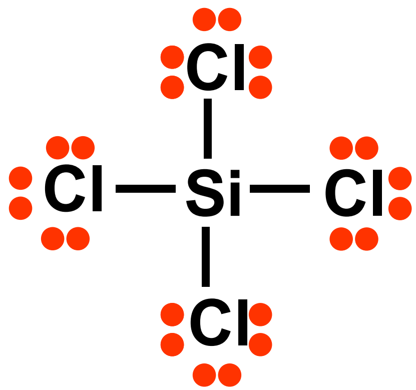
Be: Electron–groups = 2

Hybridization of Be = ***sp***

Ans: Example 2

4.3

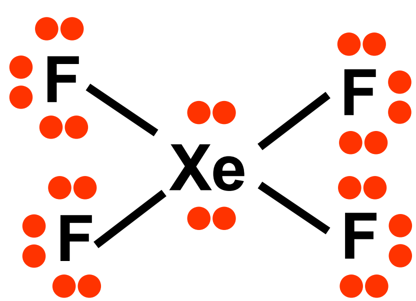
(c) $SiCl_4$ Lewis structure:



Si: Electron–groups = 4

Hybridization of Si = sp^3

(d) XeF_4 Lewis structure:

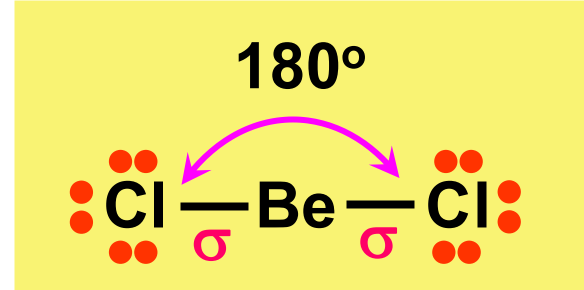


Xe: Electron–groups = 6

Hybridization of Xe = sp^3d^2

sp HYBRIDIZATION

EXAMPLE: gaseous $BeCl_2$

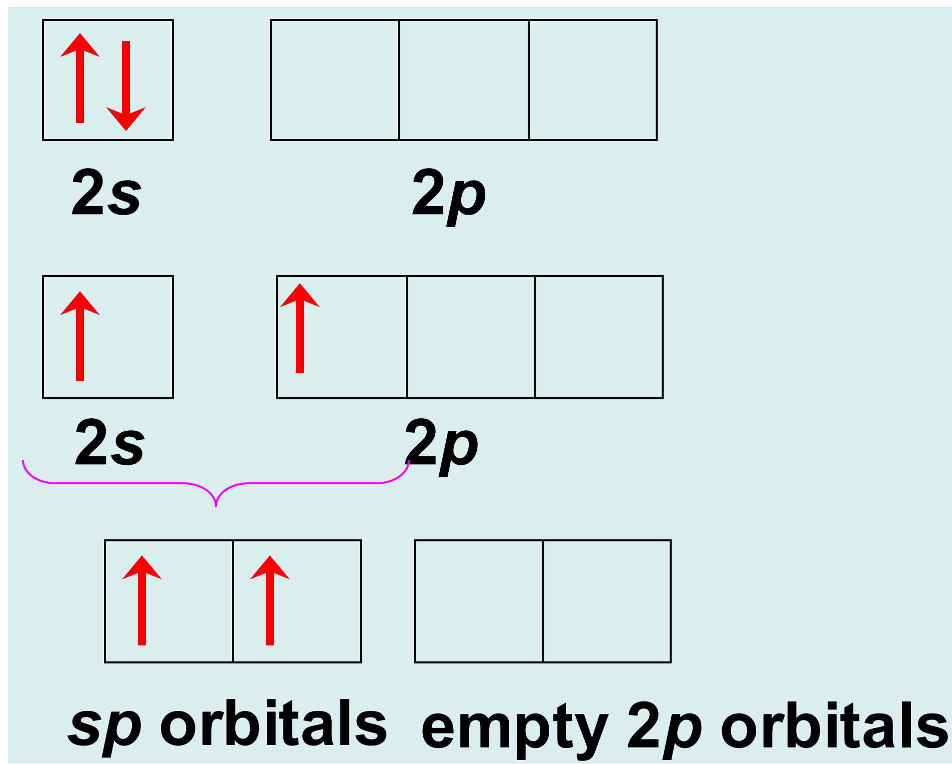


e^- configuration of Be : $1s^2\ 2s^2$

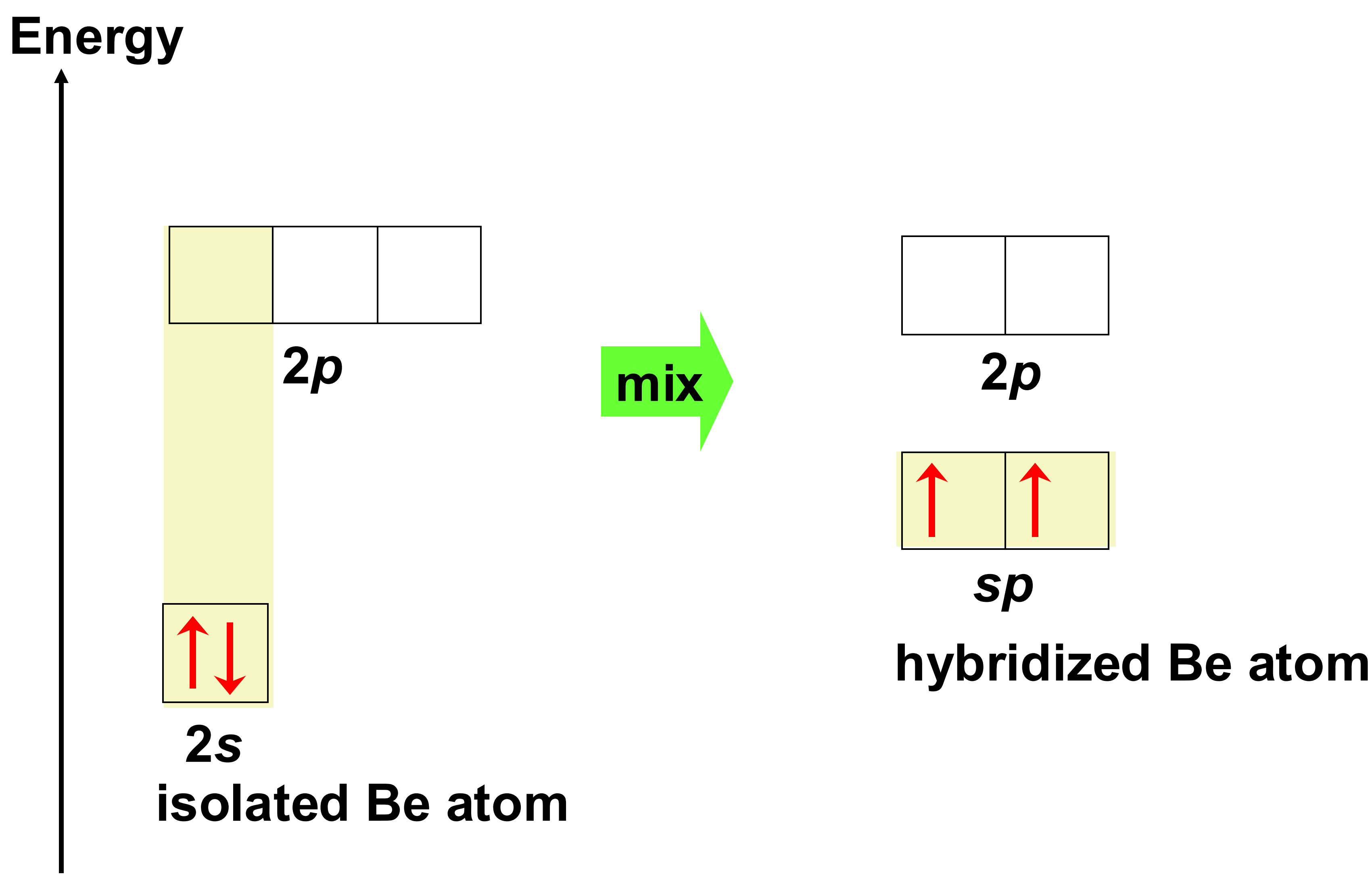
e^- configuration of Cl : $1s^2\ 2s^2\ 3s^2 3p^5$

Valence e^- in Be :

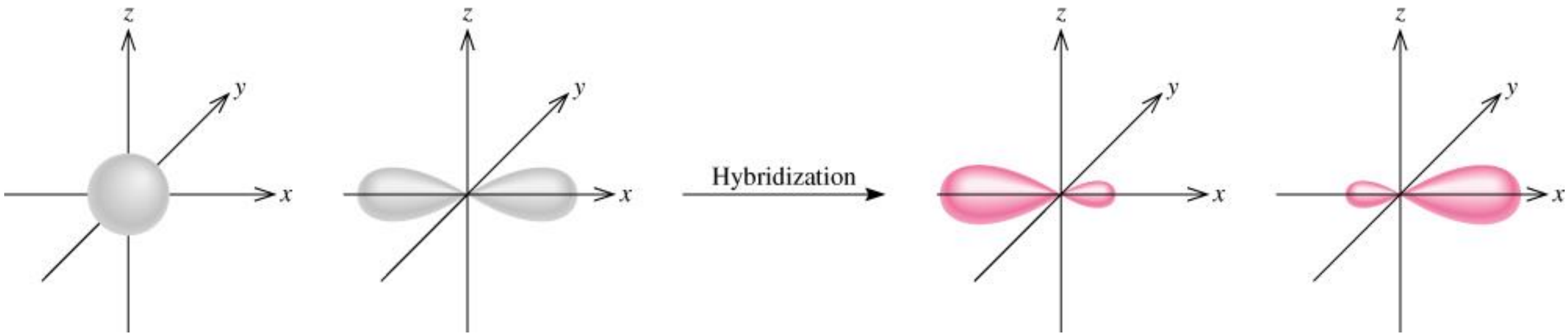
- Ground state:
- Promotion of e^-:
- *sp* hybridization:



sp hybridization



one *s* orbital + one *p* orbital → two equivalent *sp* orbitals

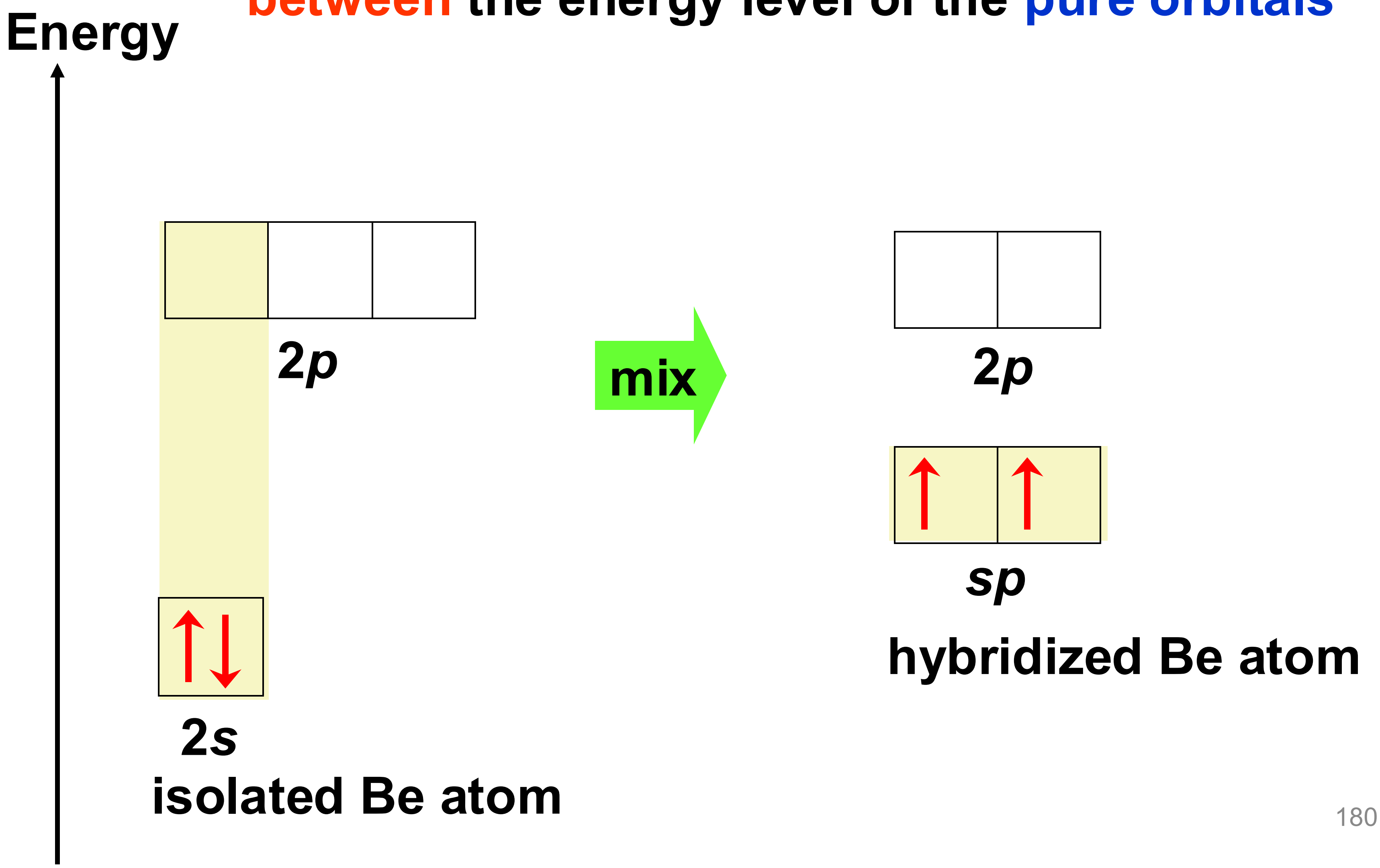


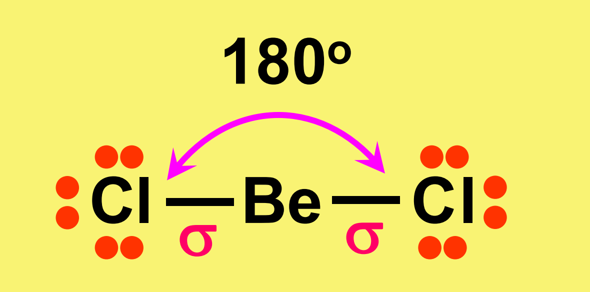
2*s* **2*p*$_x$** **two *sp* orbitals**

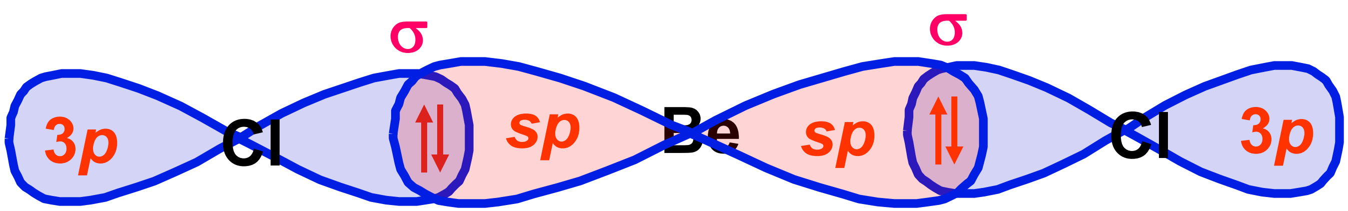
one *s* orbital + one *p* orbital → two equivalent *sp* orbitals

Keep in mind!

- Energy level of the hybrid orbitals is in between the energy level of the pure orbitals







- **Each Be–Cl (σ) bond formed by:**
  Overlap of one *sp* hybrid of Be atom and one 3*p* orbital of Cl atom

- **Two equivalent *sp* hybrid orbitals that lie 180° apart**
  ☞ **two e⁻ groups**
  (from VSEPR theory)

- **e⁻ group arrangement = linear**

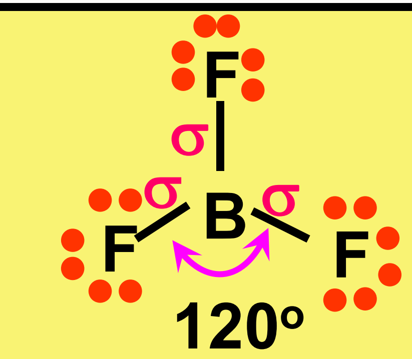
  Molecular shape = linear

sp^2 HYBRIDIZATION

EXAMPLE: BF_3

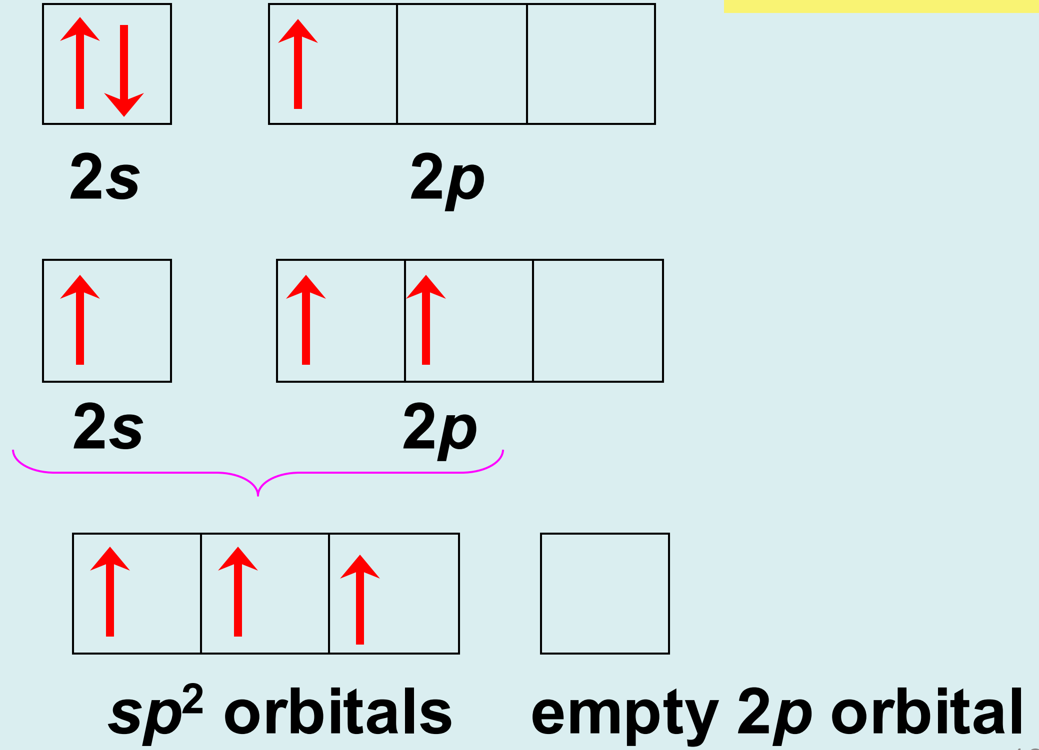
e^- configuration of B : $1s^2$ $2s^2$ $2p^1$

e^- configuration of F : $1s^2$ $2s^2$ $2p^5$

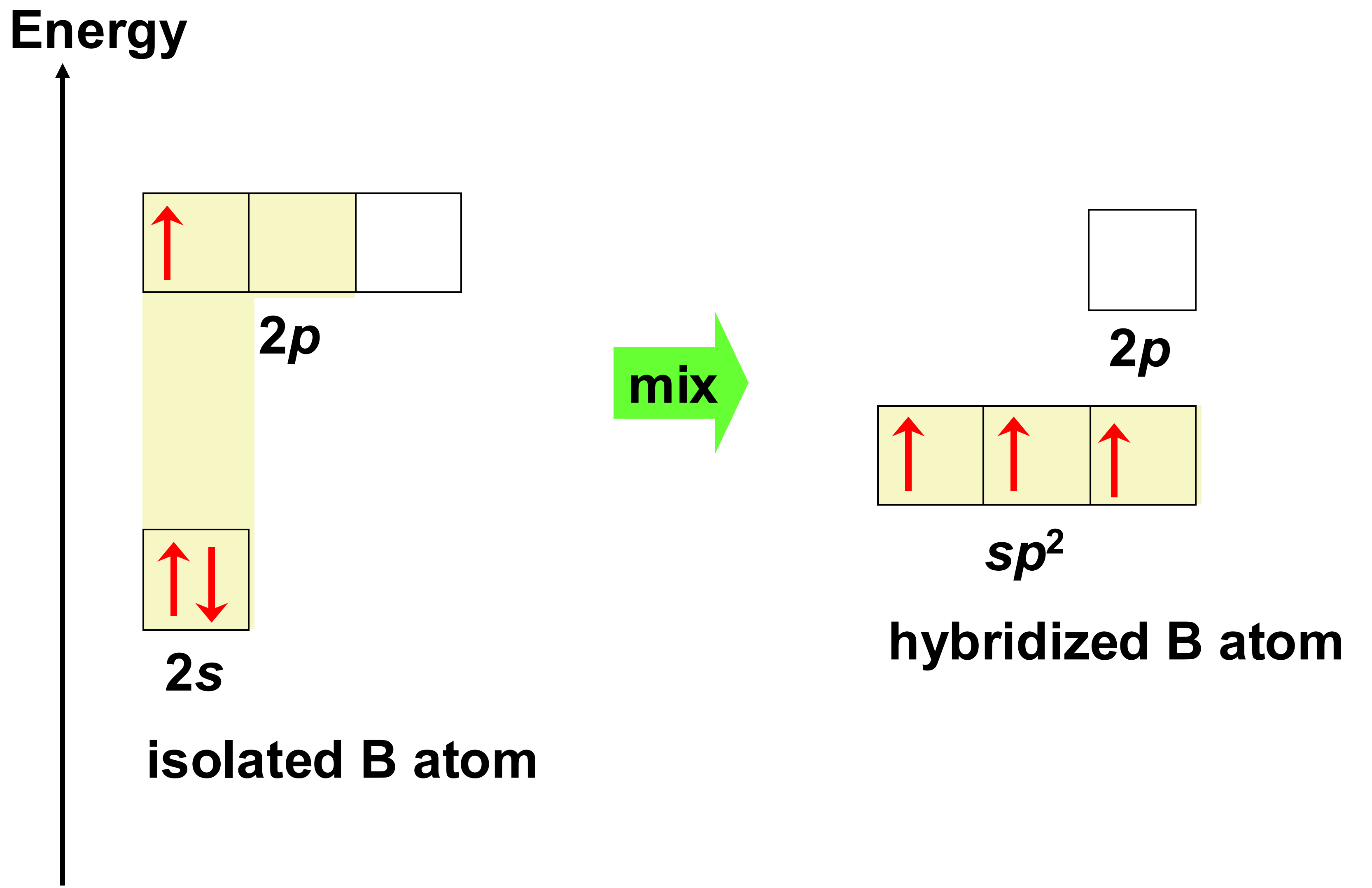


Valence e^- in B:

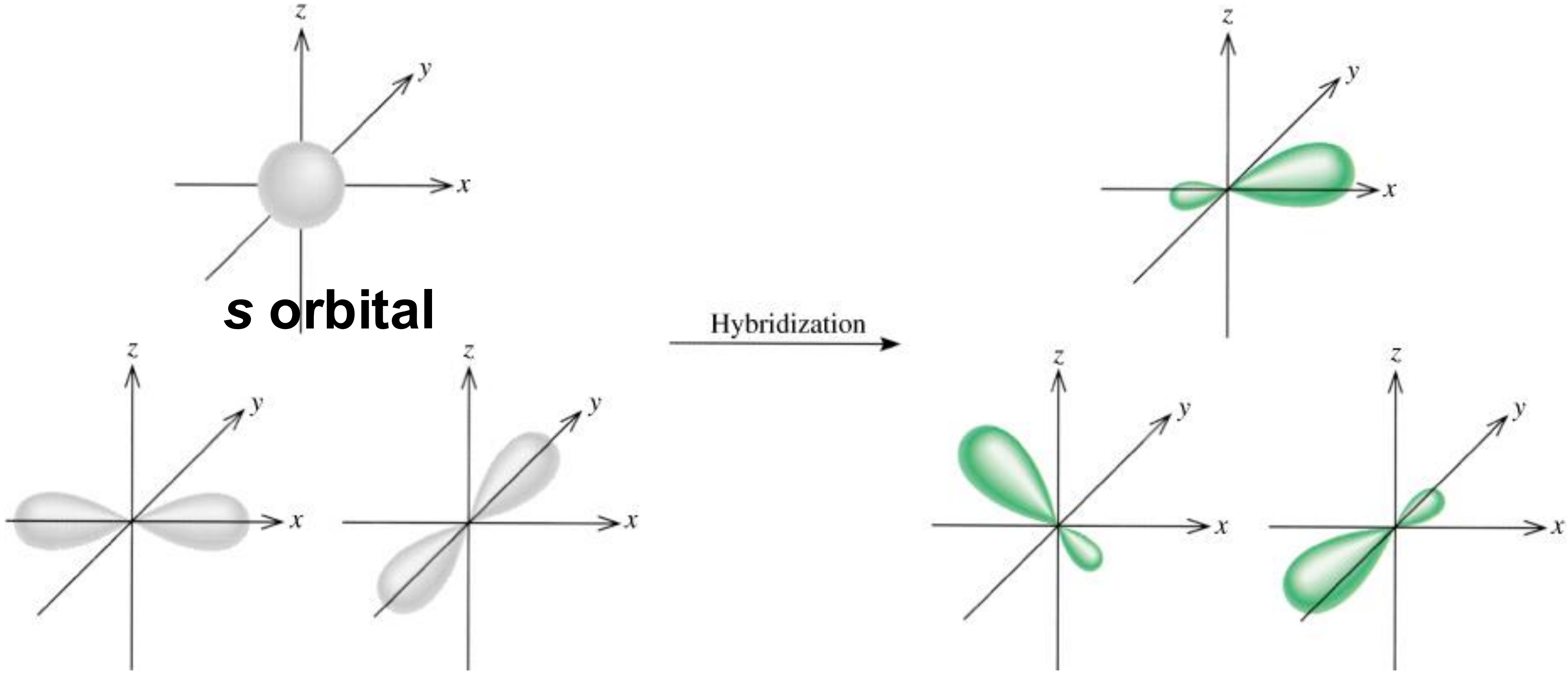
- **Ground state:**
- **Promotion of e^-:**
- **sp^2 hybridization:**



sp^2 hybridization



one *s* orbital + two *p* orbital → three equivalent *sp*^{2} orbitals

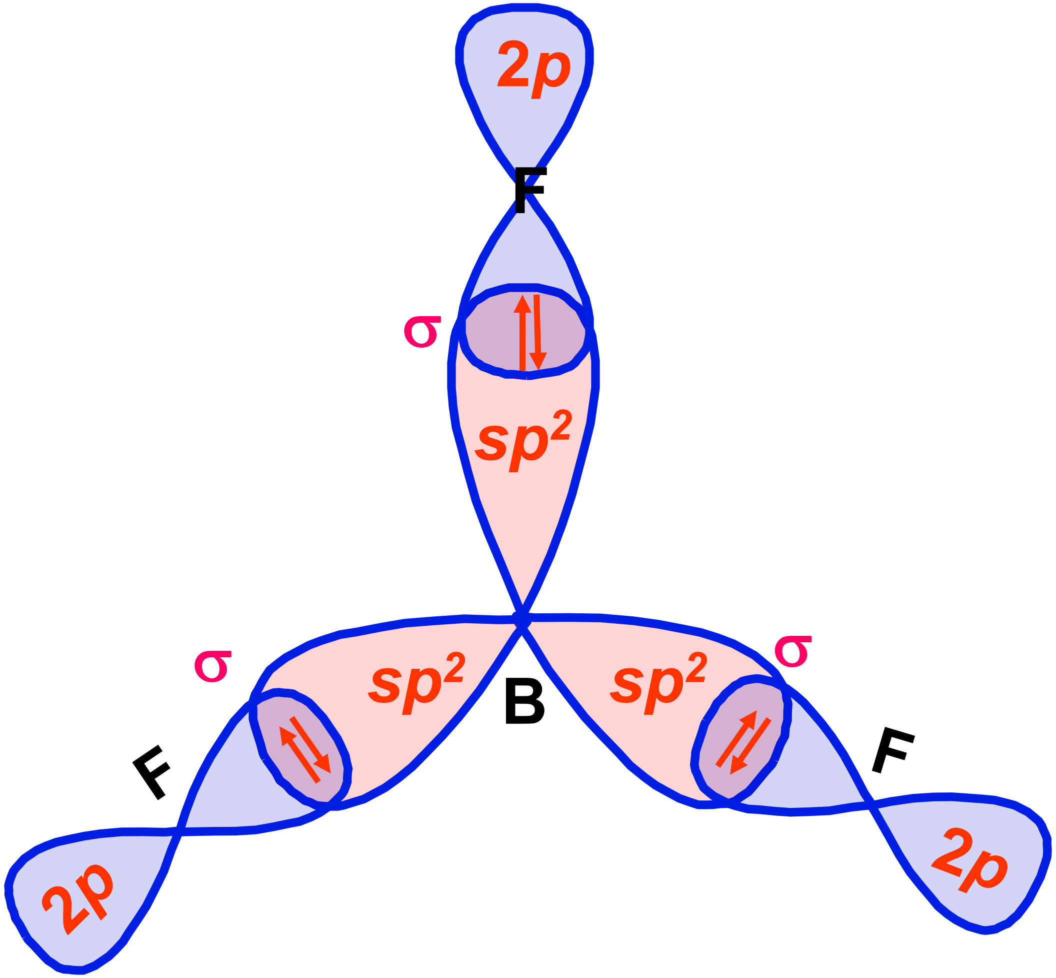


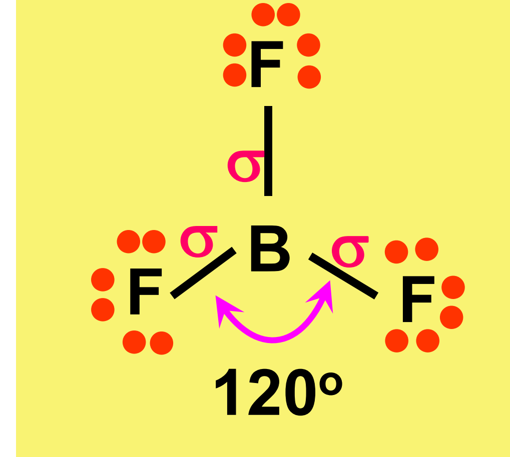
s **orbital**

p **orbitals**

three ***sp*^{2}** **orbitals**

one *s* orbital + two *p* orbital → three equivalent *sp*^{2} orbitals





- Each B–F (σ) bond formed by: Overlap of one sp^2 hybrid of B atom and one $2p$ orbital of F atom

- **Three equivalent *sp*^{2} hybrid orbitals that lie 120^o apart**
  ☞ **three e$^-$ groups**
  (from VSEPR theory)

- **e$^-$ group arrangement = trigonal planar**

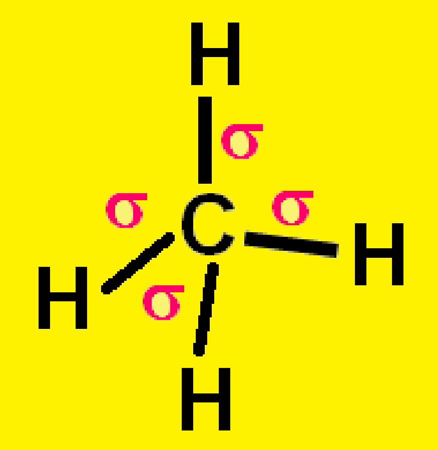
  Molecular shape = trigonal planar

sp^3 HYBRIDIZATION

EXAMPLE: CH_4

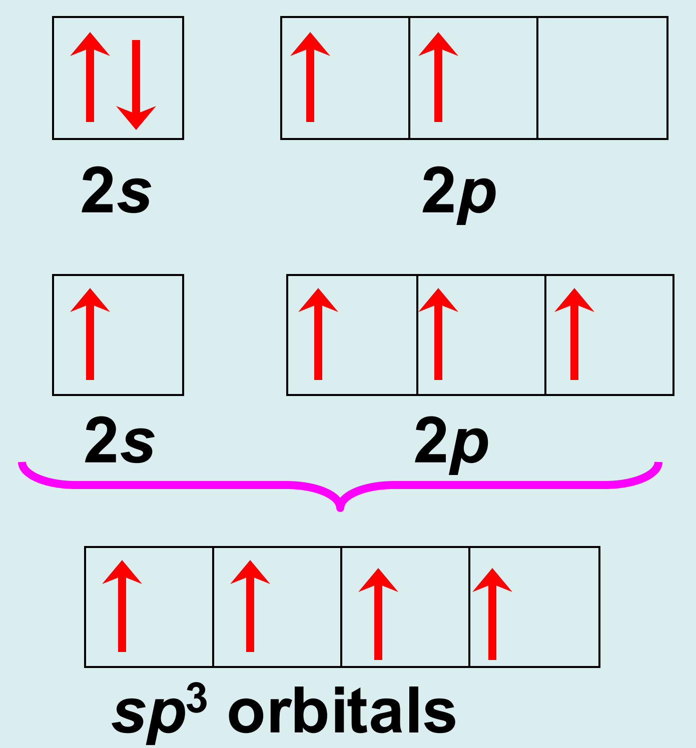
e^- configuration of C : $1s^2\ 2s^2\ 2p^2$

e^- configuration of H : $1s^1$

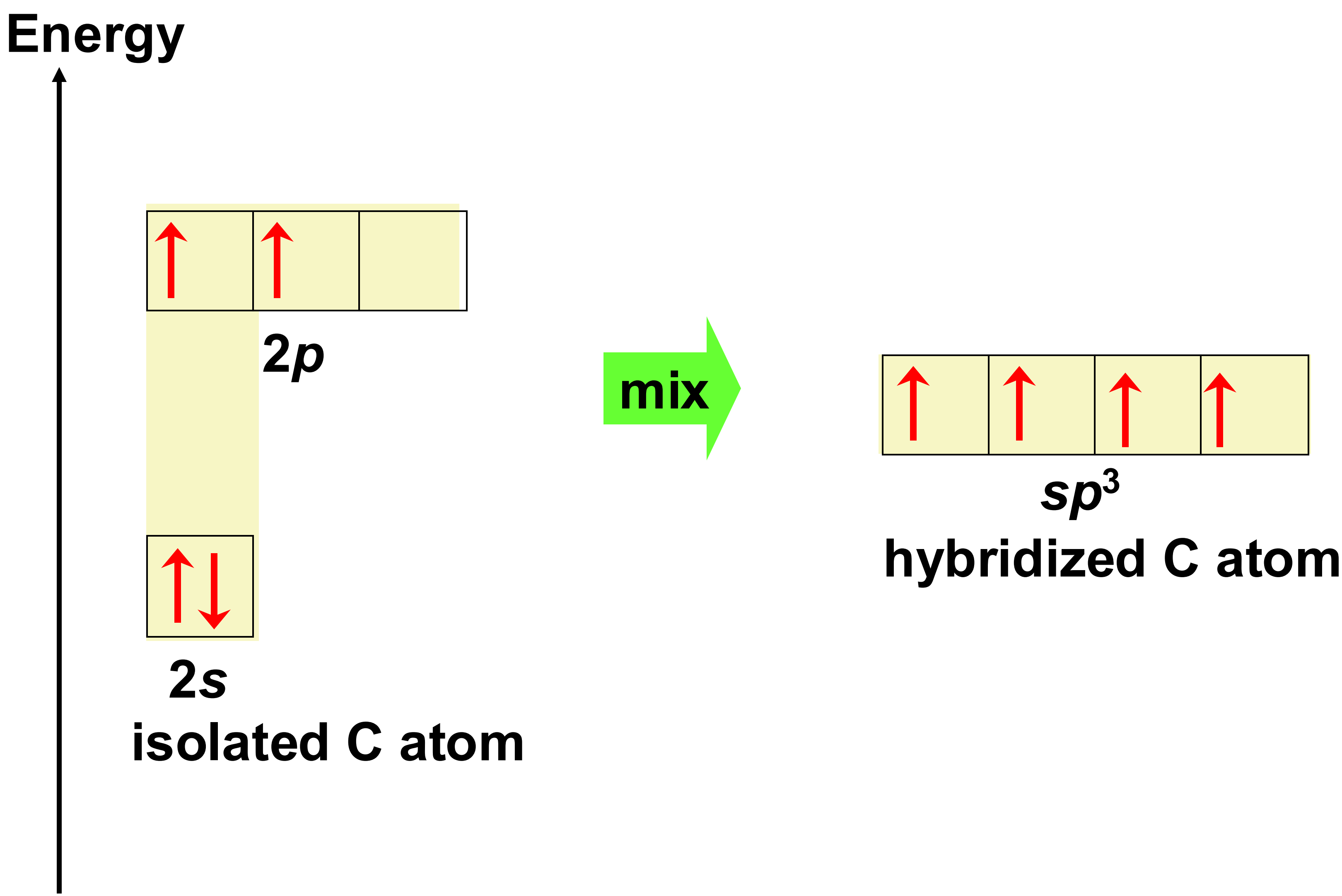


Valence e^- in C:

- Ground state:
- Promotion of e^-:
- sp^3 hybridization:

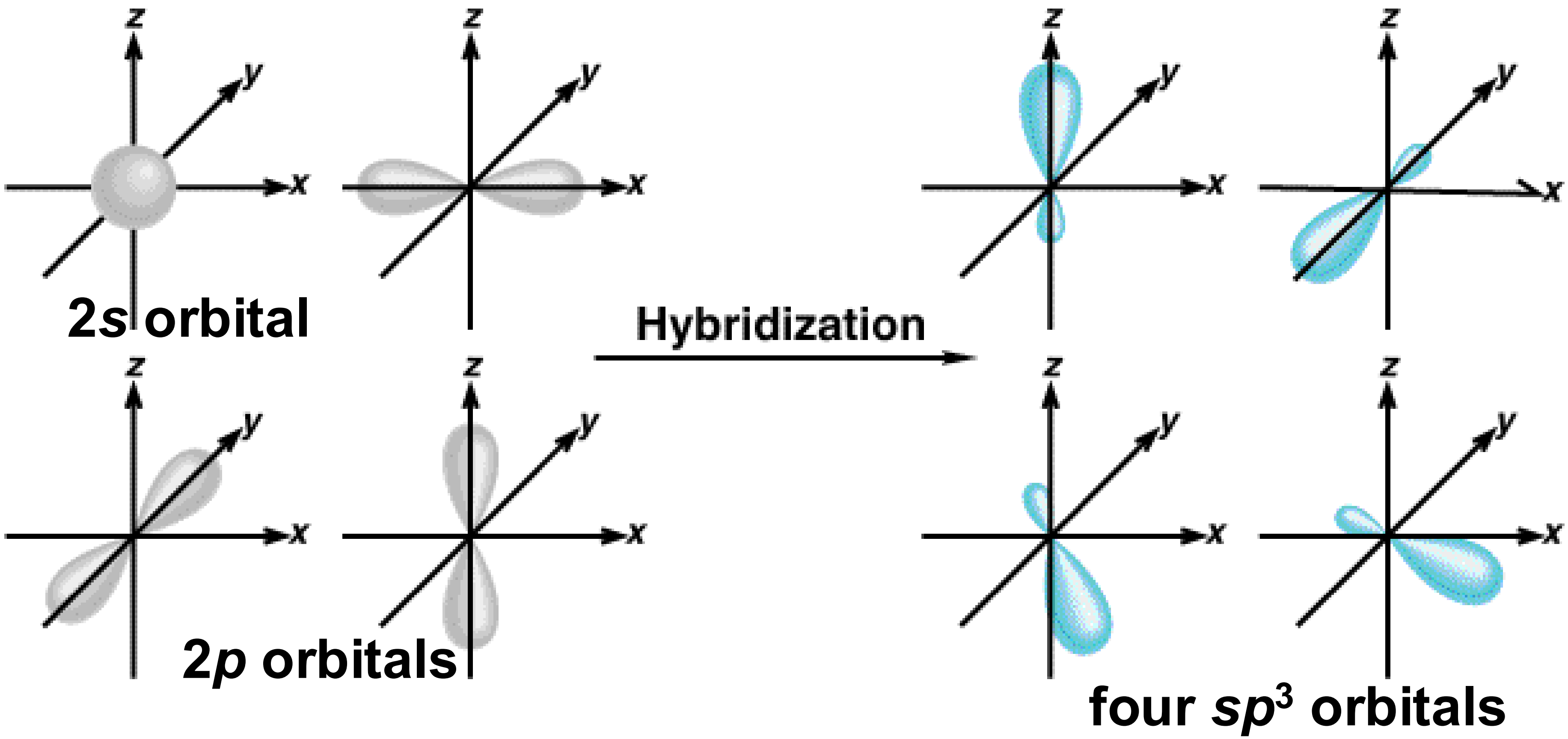


sp^3 hybridization

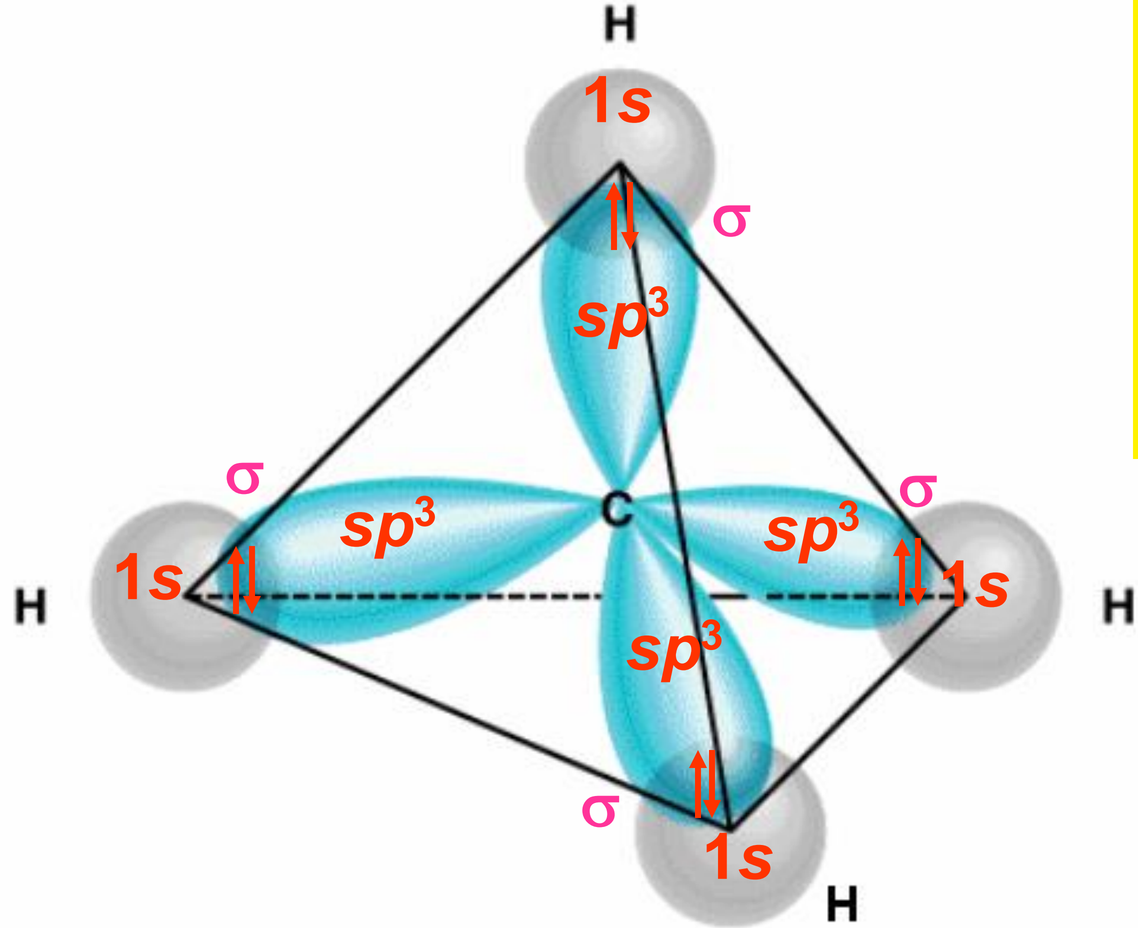


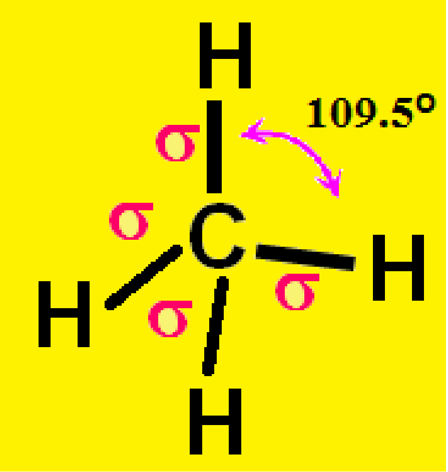
one *s* orbital + three *p* orbital → four equivalent *sp*^{3} orbitals

Formation of sp^3 Hybrid Orbitals



one *s* orbital + three *p* orbital → four equivalent sp^3 orbitals





- Each C–H (σ) bond formed by: Overlap of **one** sp^3 hybrid of C atom and **one** 1*s* orbital of H atom

- ❑ **Four equivalent *sp*^{3} hybrid orbitals that lie 109.5^o apart**

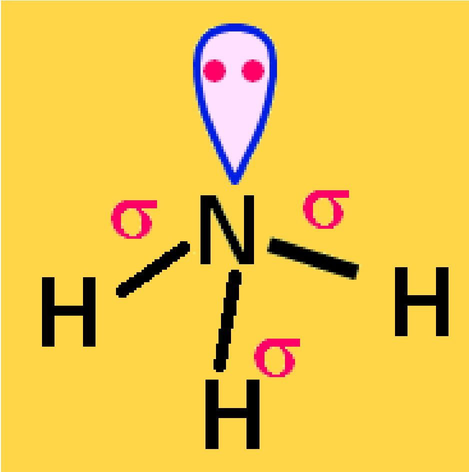
  ☞ four e$^-$ groups

  (from VSEPR theory)

- ❑ **e$^-$ group arrangement = tetrahedral**

  Molecular shape = tetrahedral

EXAMPLE: NH_3

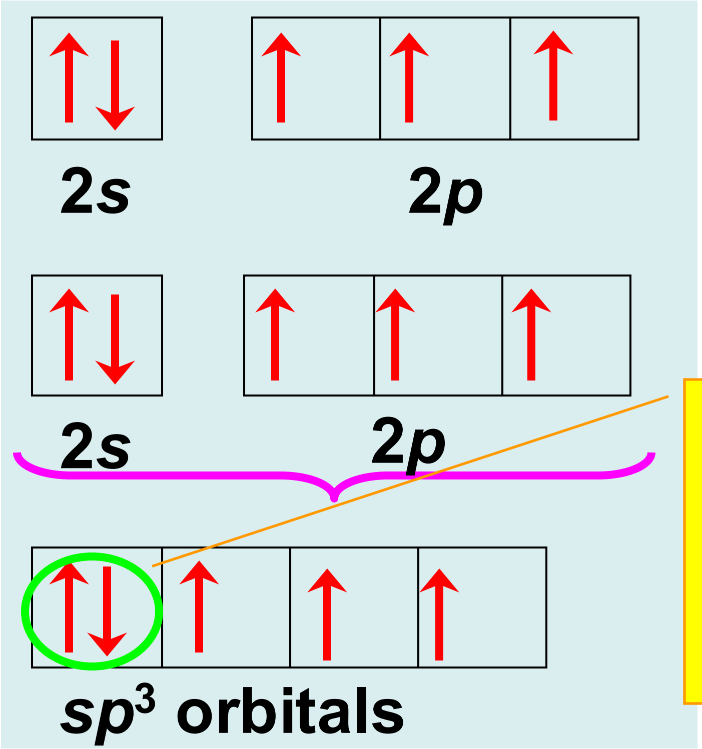


e^- configuration of N : $1s^2$ $2s^2$ $2p^3$

e^- configuration of H : $1s^1$

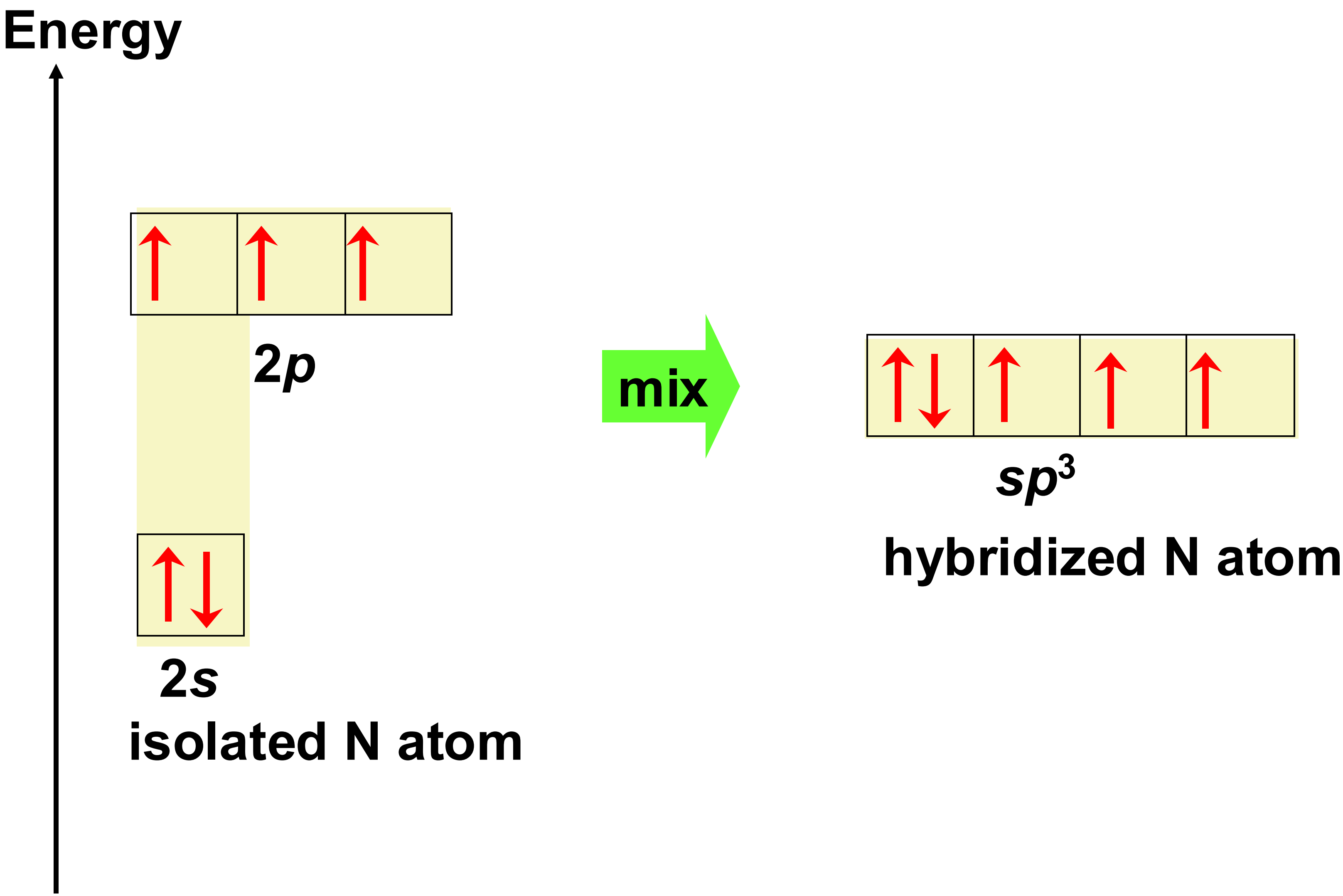
Valence e^- in N:

- Ground state:
- Promotion of e^-:
- sp^3 hybridization:



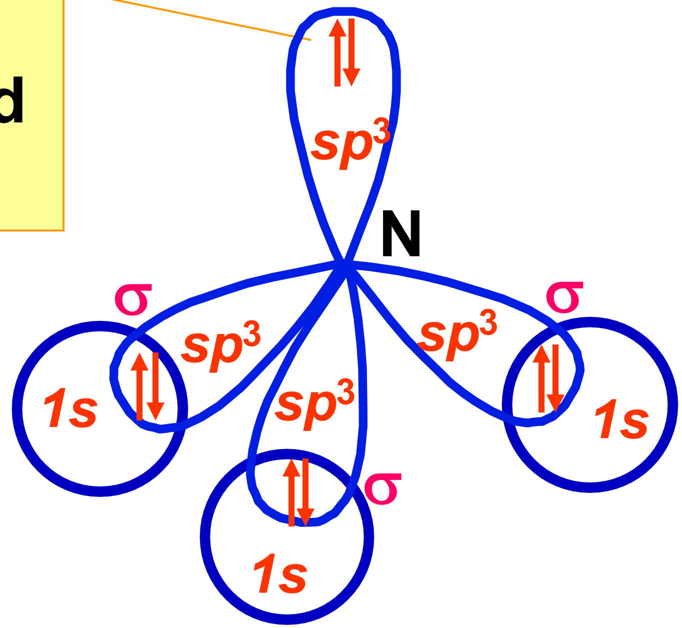
One of the sp^3 hybrid occupied by lone pair e^-

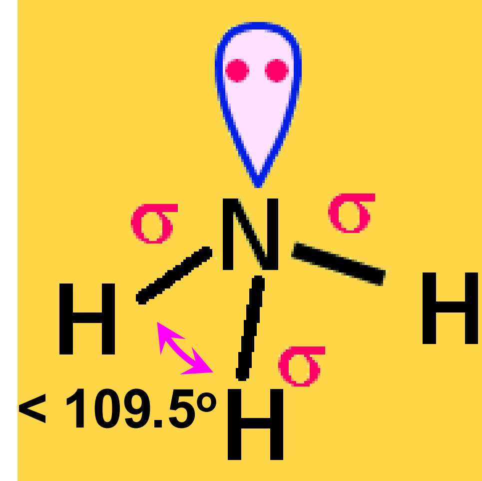
sp^3 hybridization



one *s* orbital + three *p* orbital → four equivalent *sp*^{3} orbitals

One of the sp^3 hybrid occupied by lone pair e^-





- Each **N–H** (σ) bond formed by: Overlap of **one sp^3** hybrid of N atom and **one 1*s*** orbital of H atom

- ❑ **Four *sp*^{3} hybrid orbitals that lie apart:**

  Bond angle < 109.5^o

  ☞ four e$^-$ groups

  (from VSEPR theory)

- ❑ **e$^-$ group arrangement = tetrahedral**

  Molecular shape = trigonal pyramidal

Example 3

4.3

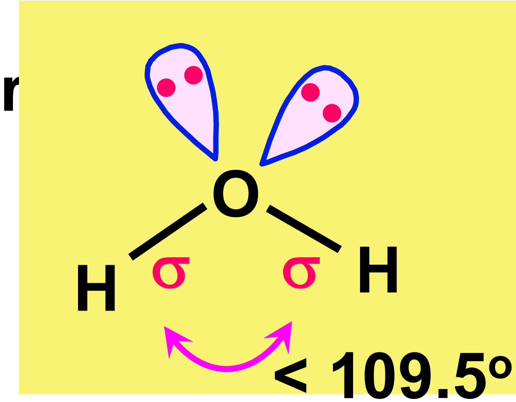
Describe a hybridization scheme for the central O atom in the molecule H_2O that is consistent with the molecular shape.

Ans: Example 3

e^- configuration of O : $1s^2$ $2s^2$ $2p^4$

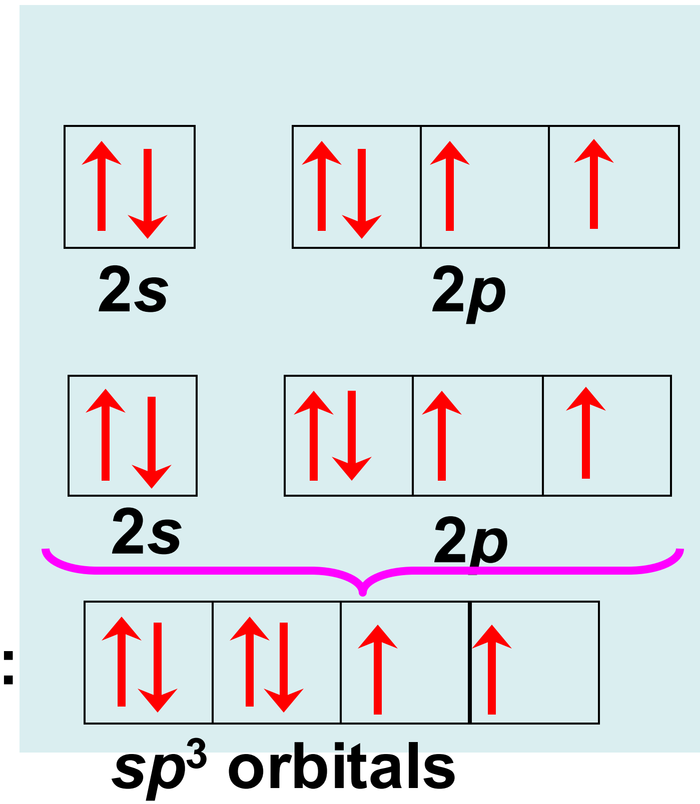
e^- configuration of H : $1s^1$

- Molecular shape:



O atom:

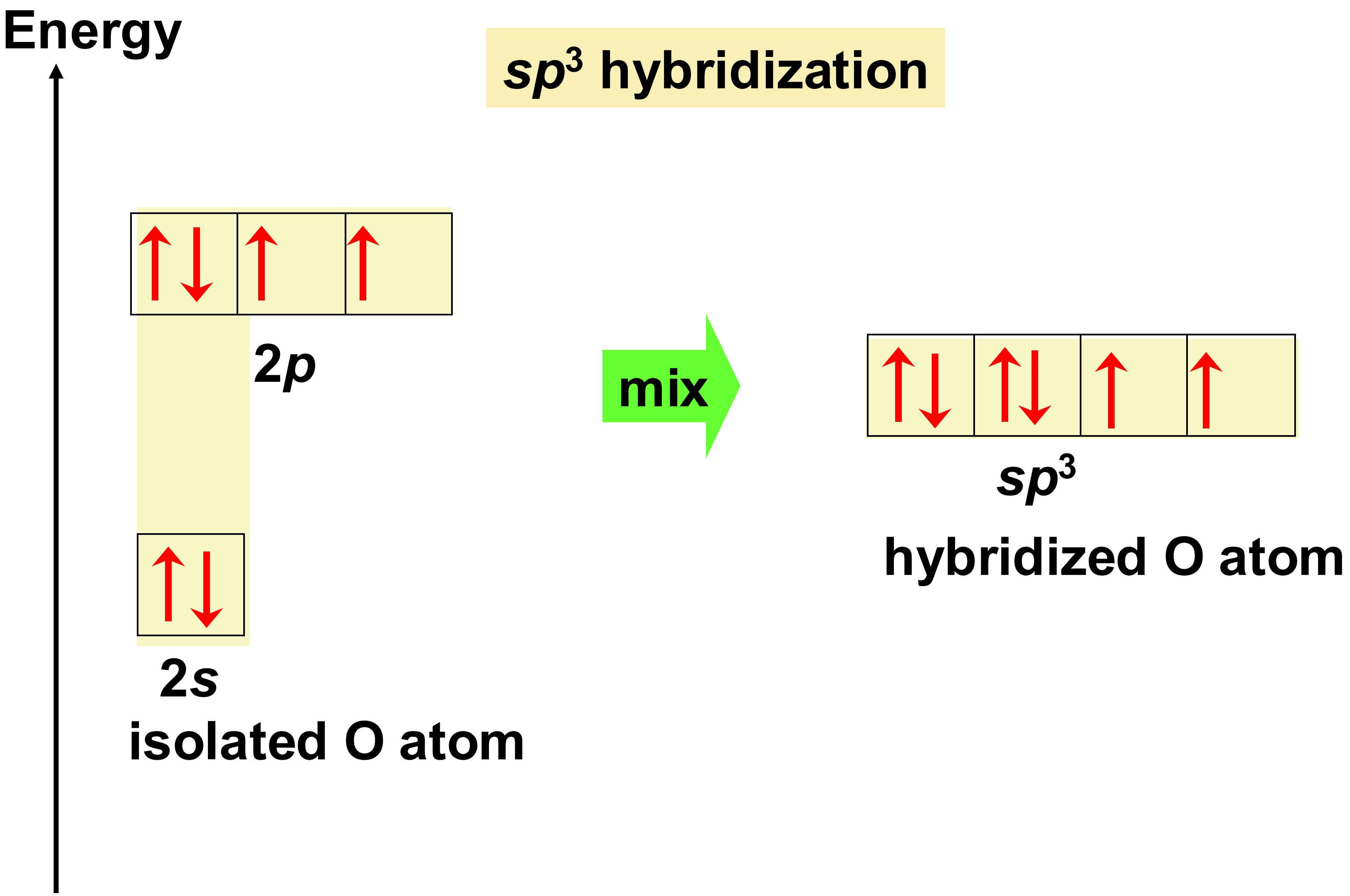
- Ground state:
- Promotion of e^-:
- sp^3 hybridization:



Ans: Example 3

4.3

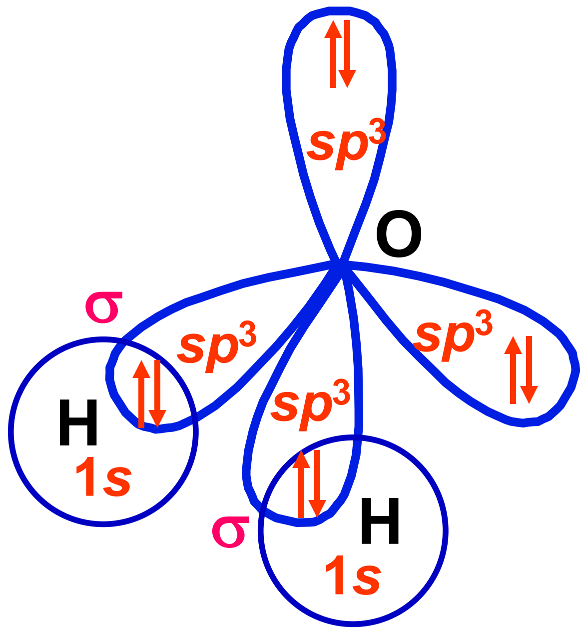
sp^3 hybridization

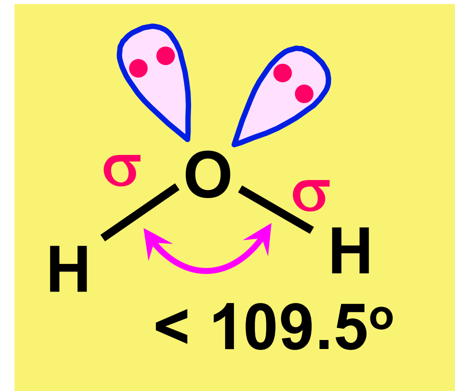


one *s* orbital + three *p* orbital → four equivalent sp^3 orbitals

Ans: Example 3

4.3





- Electron group arrangement: tetrahedral

- Molecular shape: bent (V–shaped)

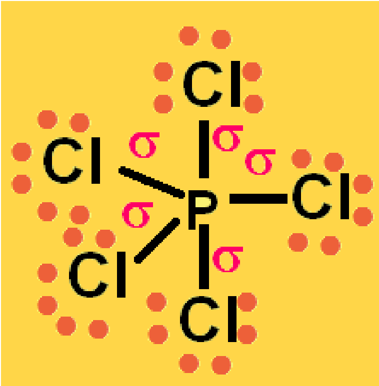
sp^3*d* HYBRIDIZATION

EXAMPLE: PCl_5

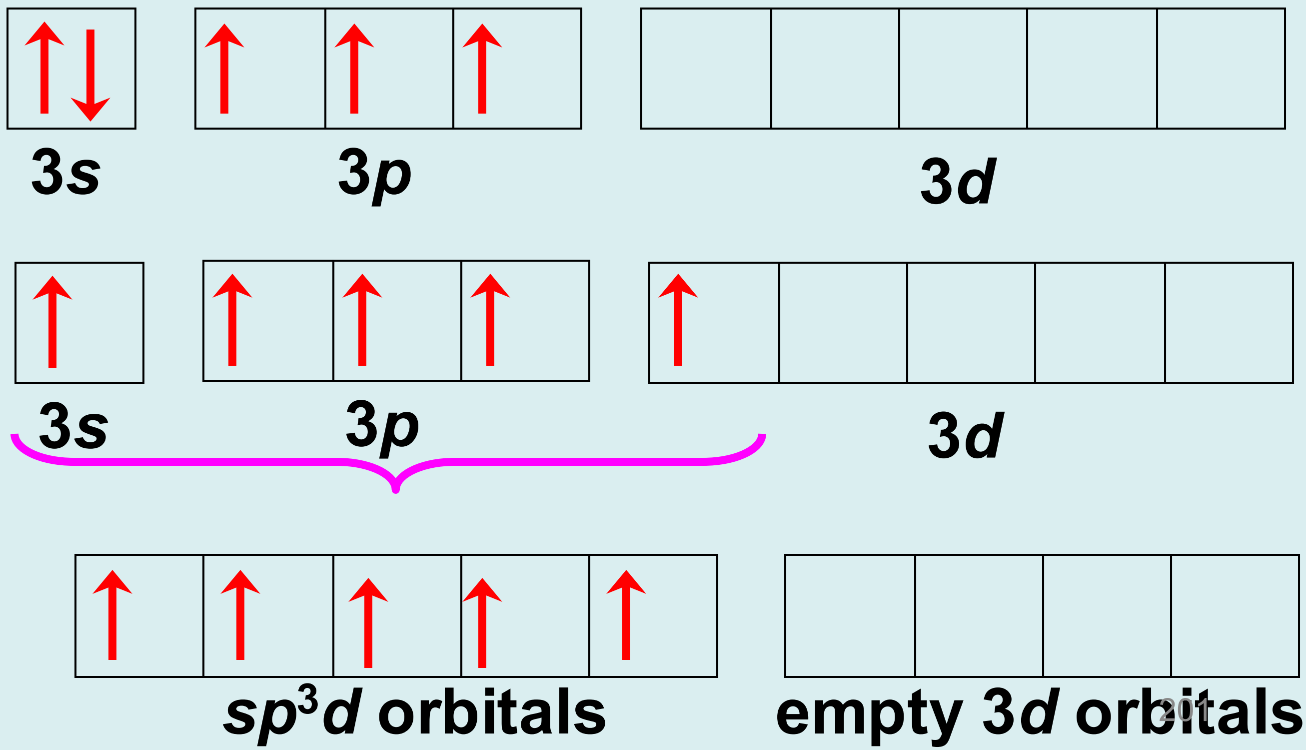
e^- configuration of P : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

e^- configuration of Cl : $1s^2\ 2s^2\ 3s^2 3p^5$

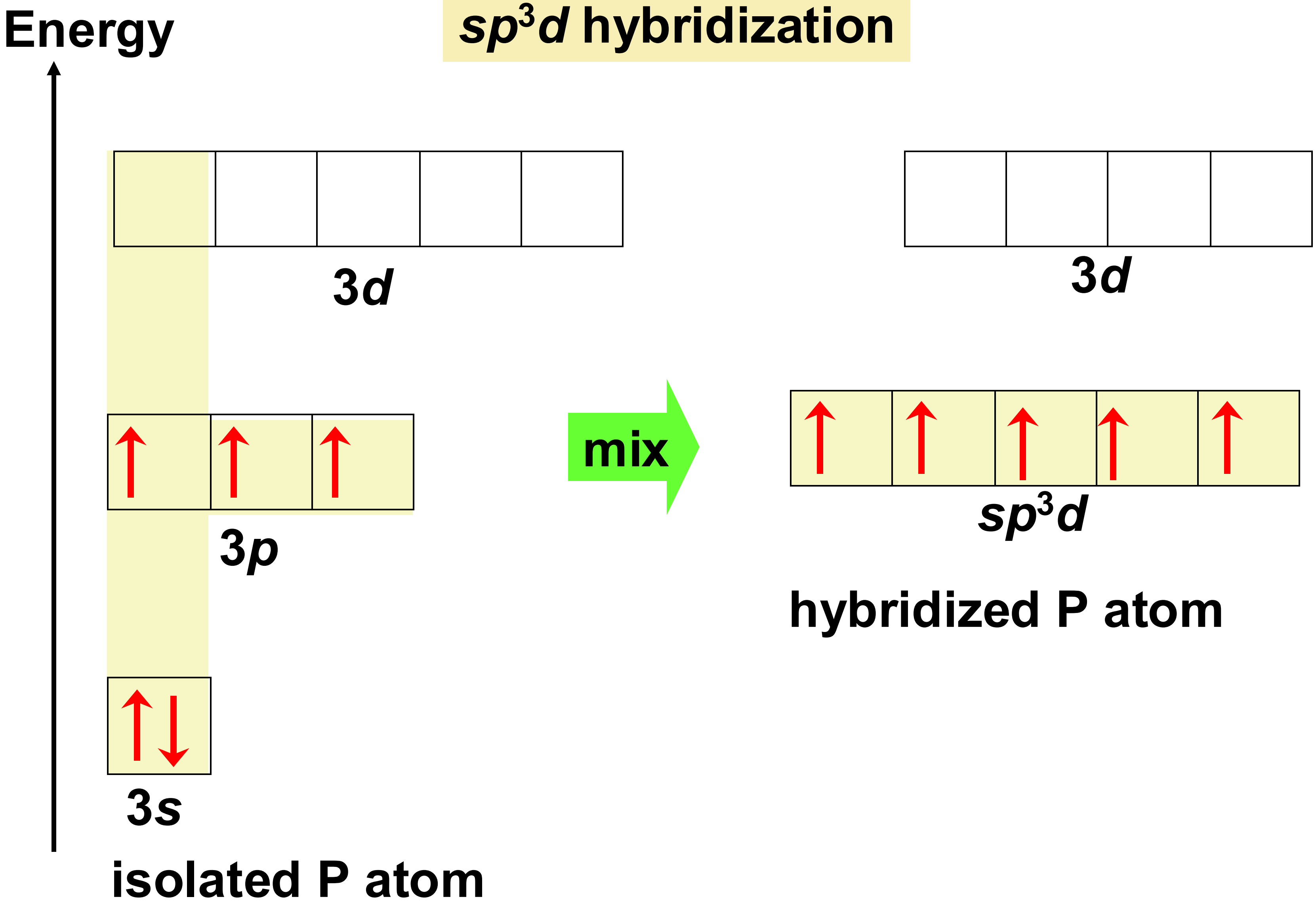
Valence e^- in P:



- Ground state:
- Promotion of e^-:
- *sp*^{3}*d* hybridization:

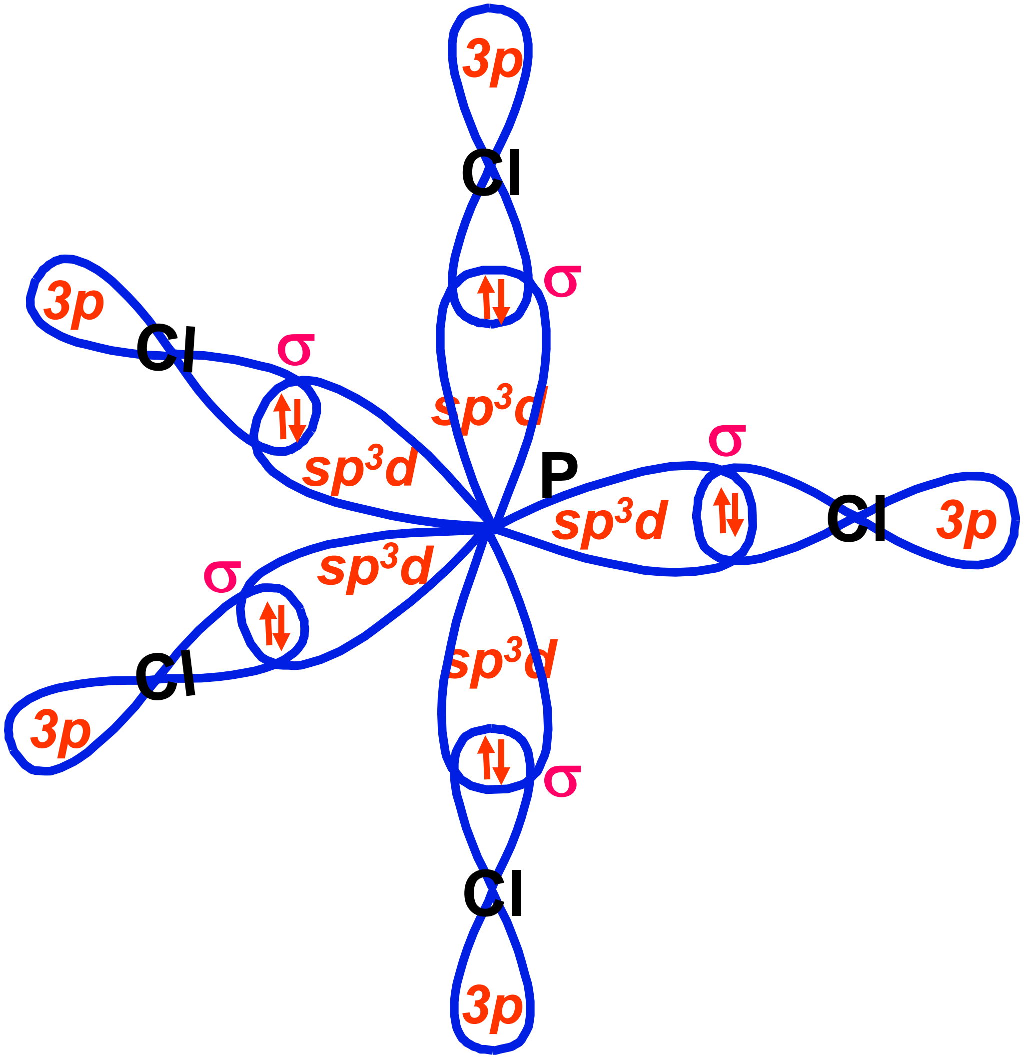


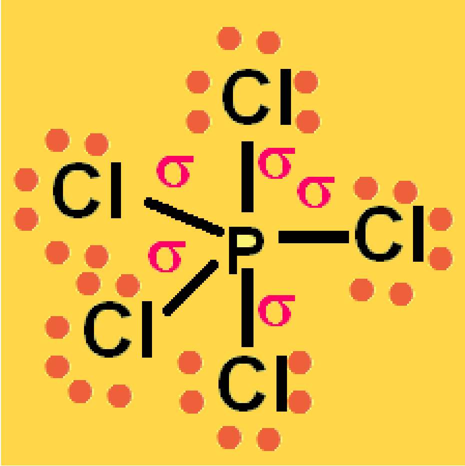
sp^3d hybridization

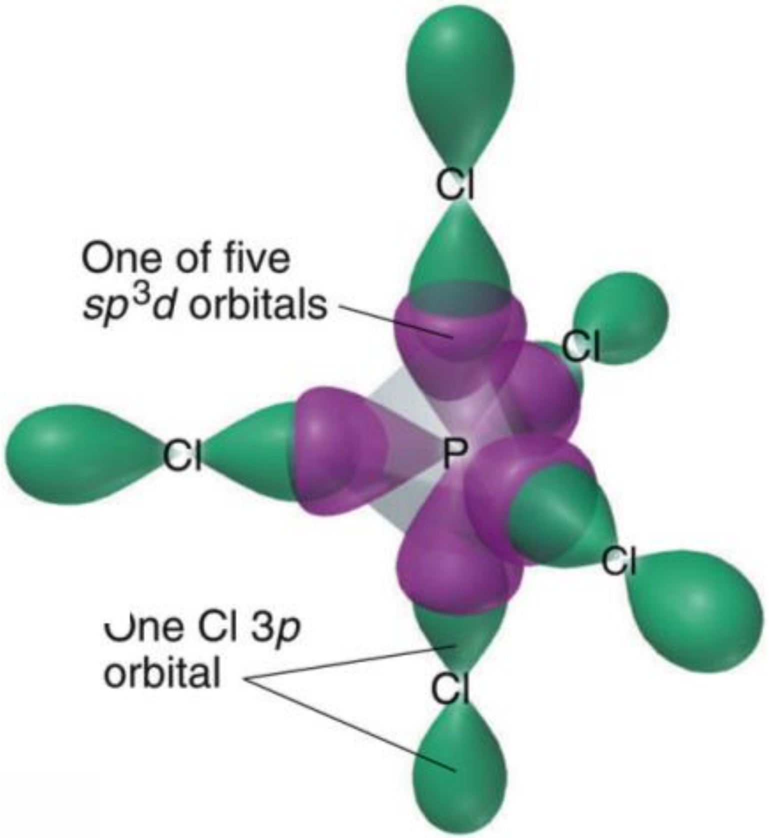


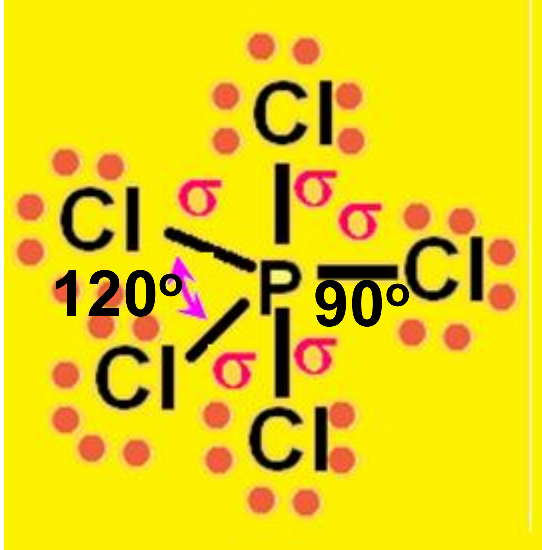
one *s* orbital + three *p* orbital + one *d* orbital

→ five equivalent sp^3d orbitals


3p
Cl
σ
sp3d
P


Cl
σ
P





- ❑ **Each P–Cl bond formed by: Overlap of one sp^3d hybrid of P atom and one 3*p* orbital of Cl atom**

- Five equivalent sp^3d hybrid orbitals that lie apart: Bond angle = 90° and 120°
  ☞ five e^- groups
  (from VSEPR theory)

- e^- group arrangement = trigonal bipyramidal

  Molecular shape = trigonal bipyramidal

Example 4

4.3

Describe a hybridization scheme for the central S atom in the molecule SF_4 that is consistent with the molecular shape.
Which orbitals of S atom are involved in overlaps, and which are by lone–pair electrons?

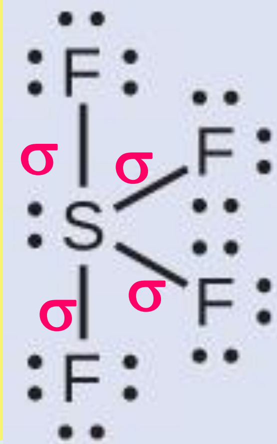
Ans: Example 4

4.3

e⁻ configuration of S : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

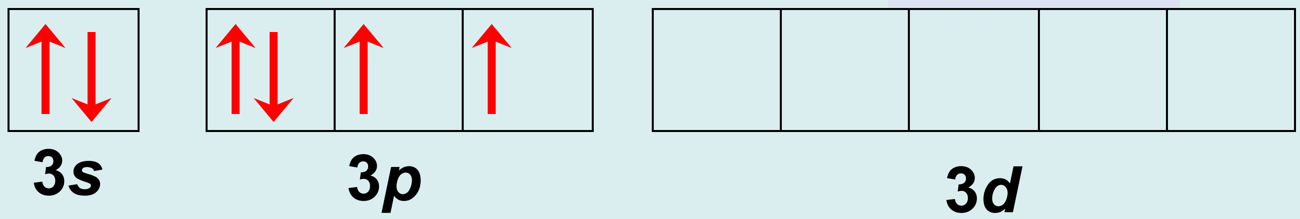
e⁻ configuration of F : $1s^2\ 2s^2\ 2p^5$

- Molecular shape:

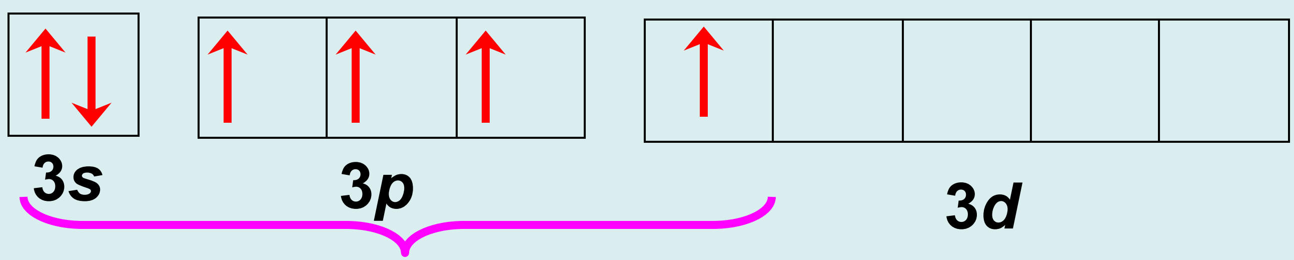


S atom:

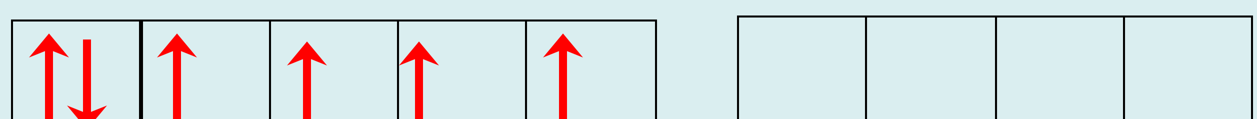
- Ground state:



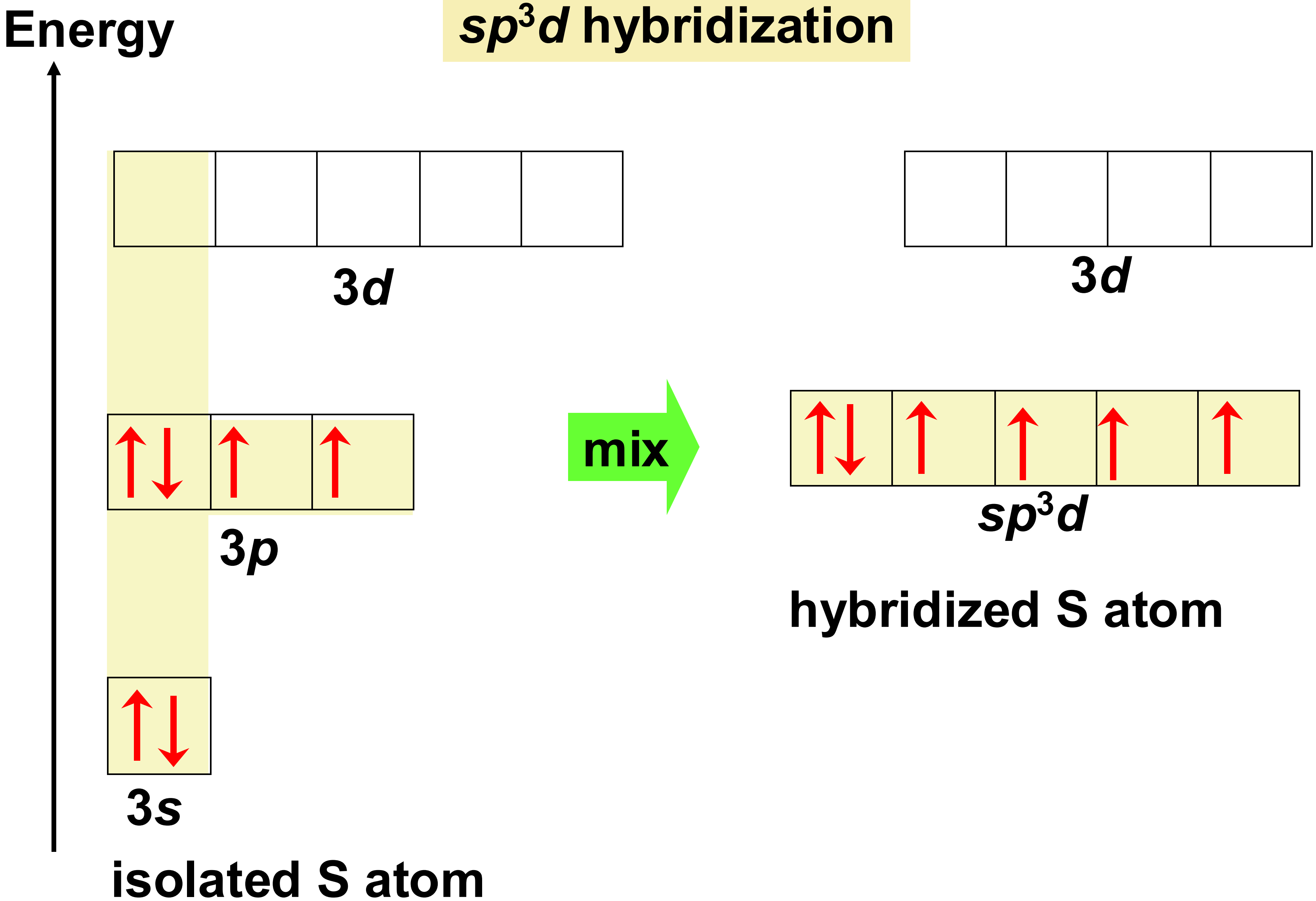
- Promotion of electron:



- sp^3d hybridization:



sp^3d hybridization

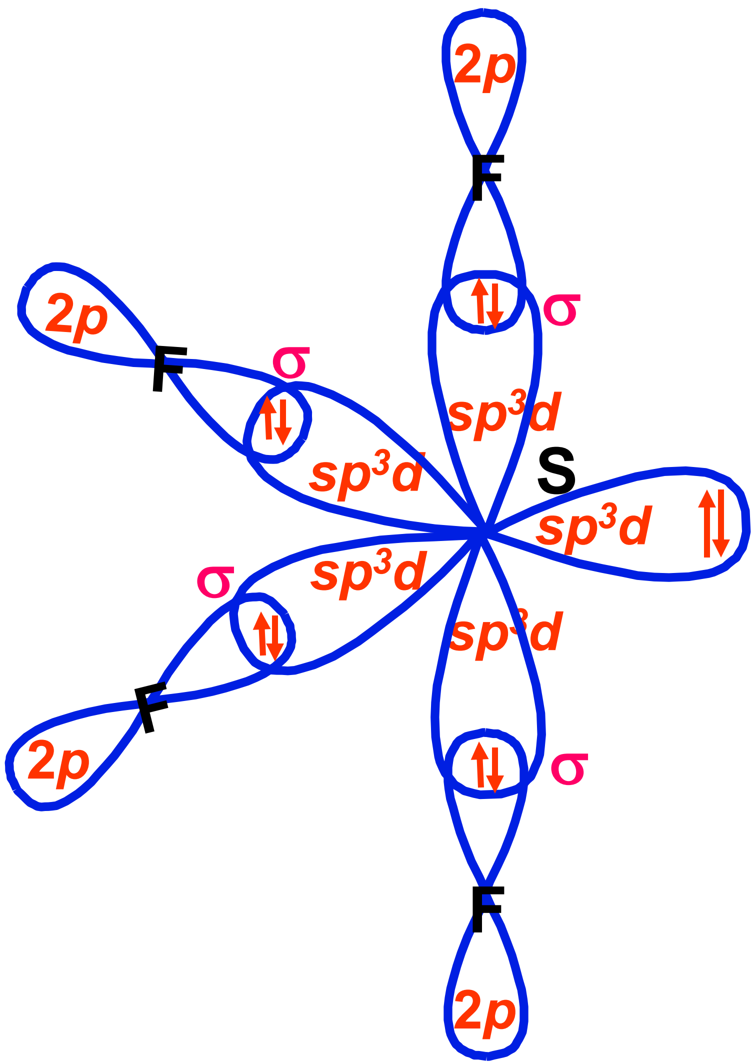


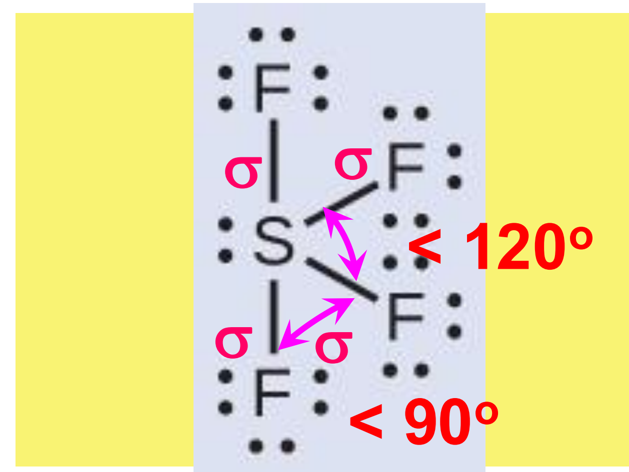
one *s* orbital + three *p* orbital + one *d* orbital

→ five equivalent sp^3d orbitals

Ans: Example 4

4.3





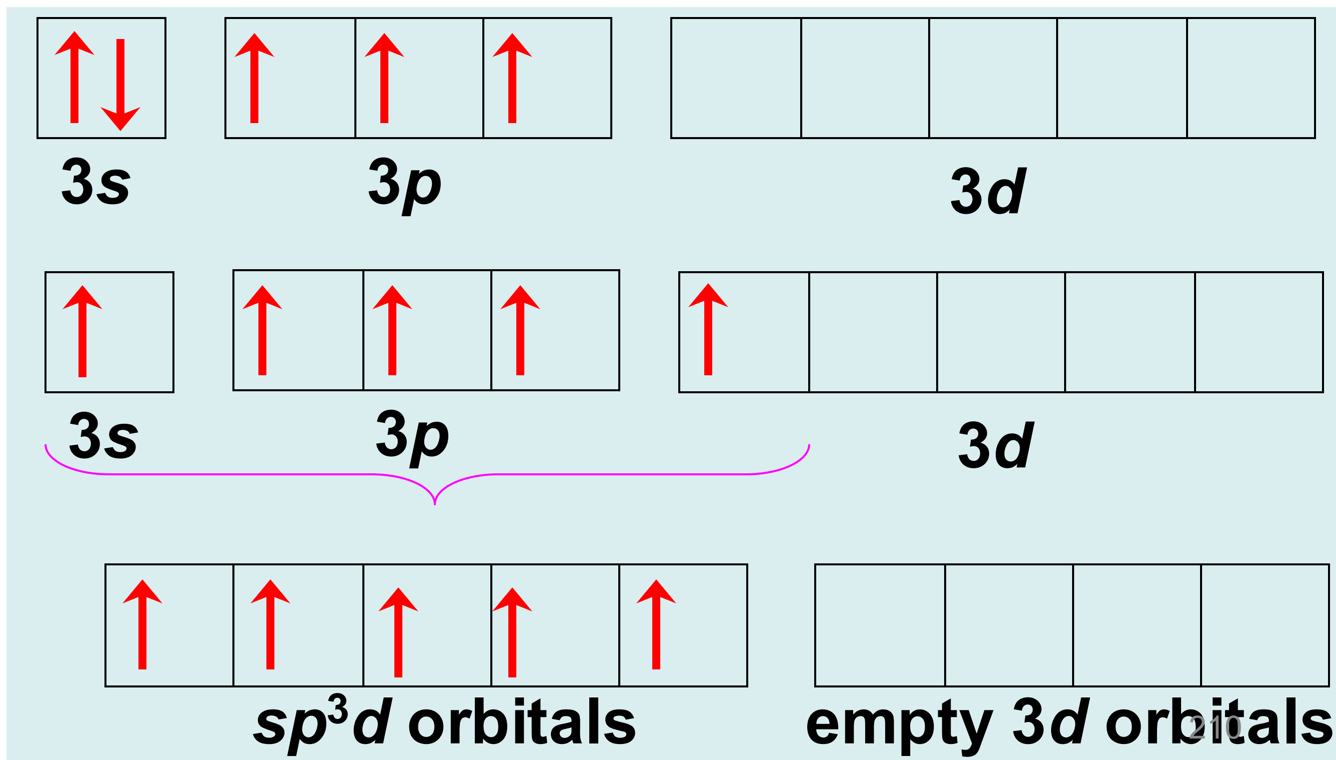
- **Electron group arrangement: trigonal bipyramidal**
- **Molecular shape: seesaw**

Keep in mind!

- Atoms of **period 3** and higher can **expand** because they have **empty *d* orbitals**

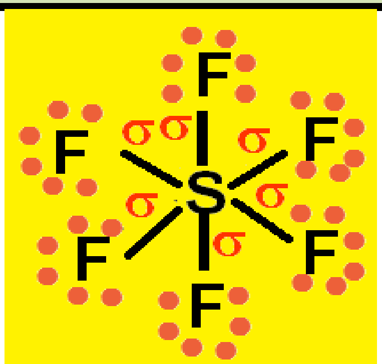
Valence e$^-$ in P:

- **Ground state:**
- **Promotion of e$^-$:**
- ***sp^{3}d* hybridization:**



sp^3d^2 HYBRIDIZATION

EXAMPLE: SF_6



e^- configuration of S : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

e^- configuration of F : $1s^2$ $2s^2\ 2p^5$

Valence e^- in S:

- **Ground state:**

3s	3p			3d				
↑↓	↑↓	↑	↑					

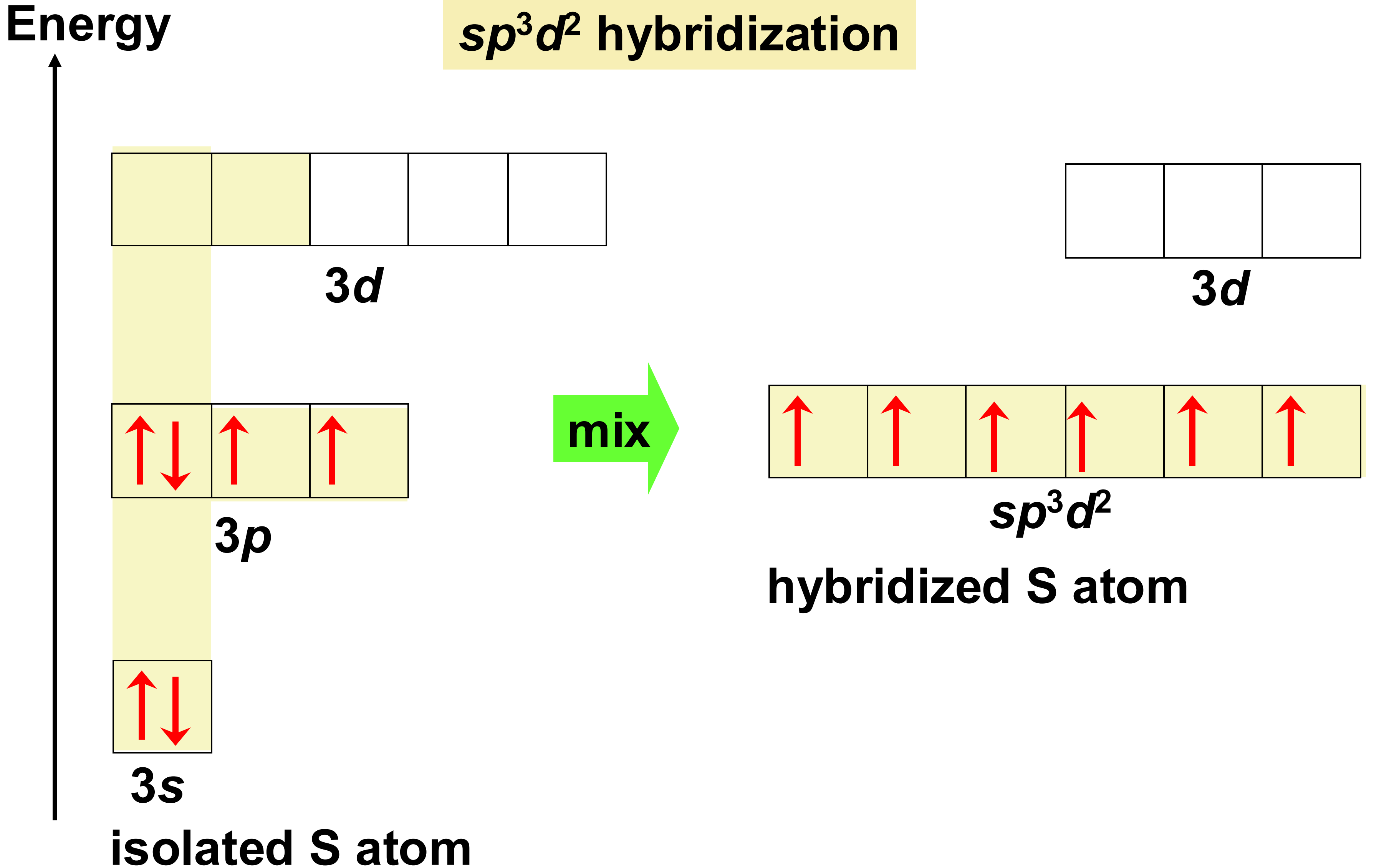
- **Promotion of e^-:**

3s	3p			3d				
↑	↑	↑	↑	↑	↑			

- ***sp^3d^2* hybridization:**

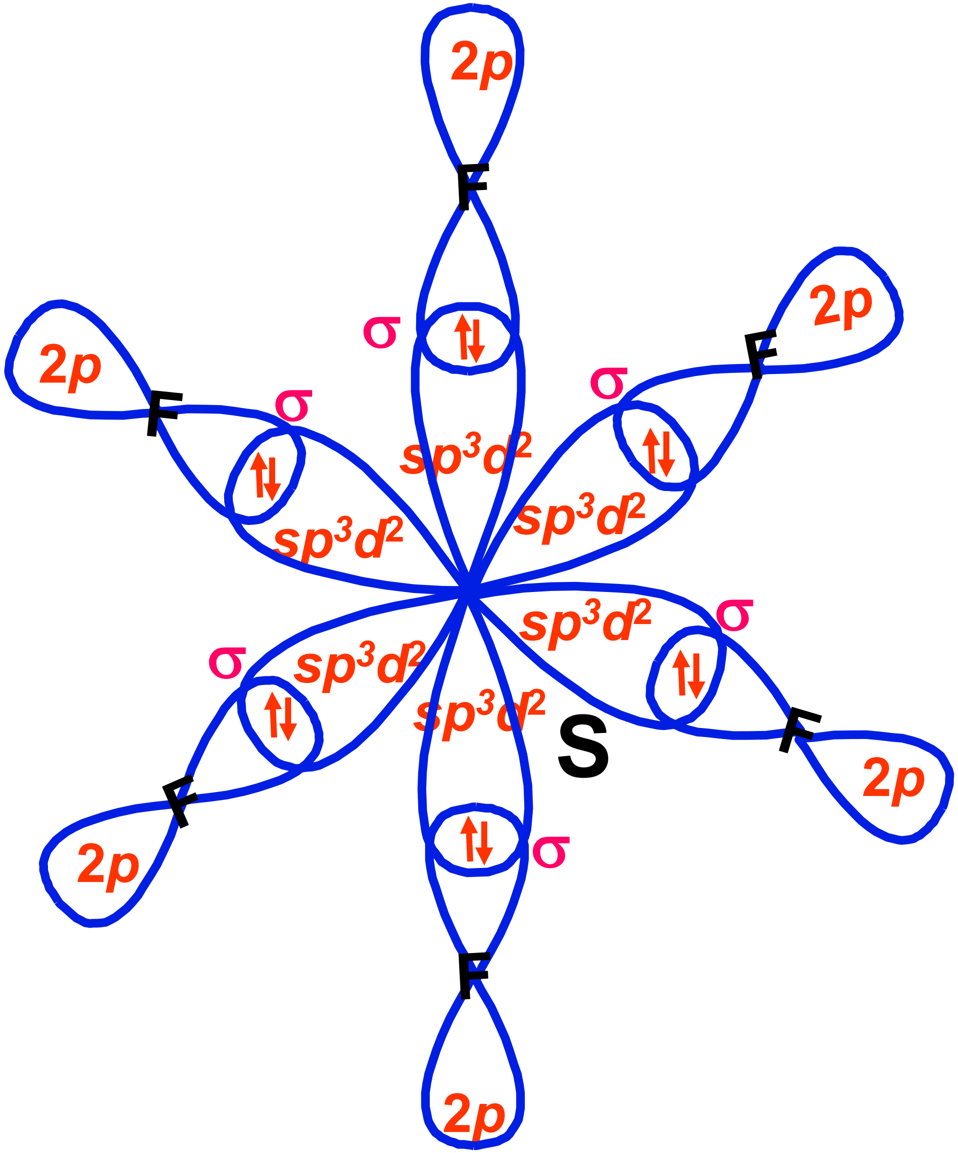
sp^3d^2 orbitals						empty 3*d* orbitals		
↑	↑	↑	↑	↑	↑			

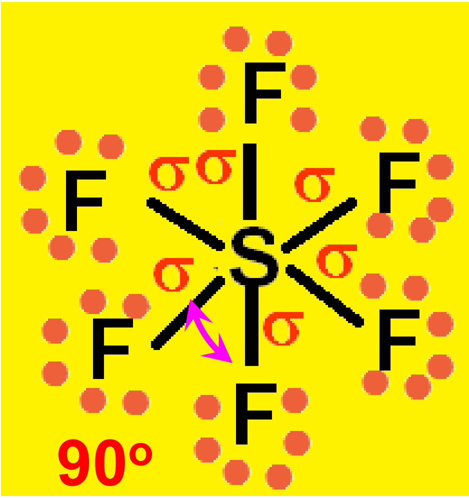
sp^3d^2 hybridization



one *s* orbital + three *p* orbital + two *d* orbitals

→ six equivalent sp^3d^2 orbitals





- ❑ **Each S–F bond formed by:**
  Overlap of one sp^3d^2 hybrid of S atom and one $2p$ orbital of F atom

- Six equivalent sp^3d^2 hybrid orbitals that lie 90° apart
  ☞ six e^- groups
  (from VSEPR theory)

- e^- group arrangement = octahedral

  Molecular shape = octahedral

Example 5

4.3

Describe a hybridization scheme for the central Xe atom in the molecule XeF_4 that is consistent with the molecular shape.
Which orbitals of Xe atom are involved in overlaps, and which are by lone–pair electrons?

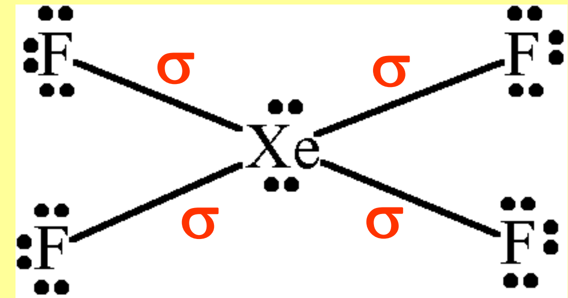
Ans: Example 5

4.3

Valence e^- configuration of Xe :

e^- configuration of F : $1s^2$ $2s^2$ $2p^5$

- Molecular shape:



Xe atom:

- Ground state: ↑↓ (5s) | ↑↓ ↑↓ ↑↓ (5p) | _ _ _ _ _ (5d)

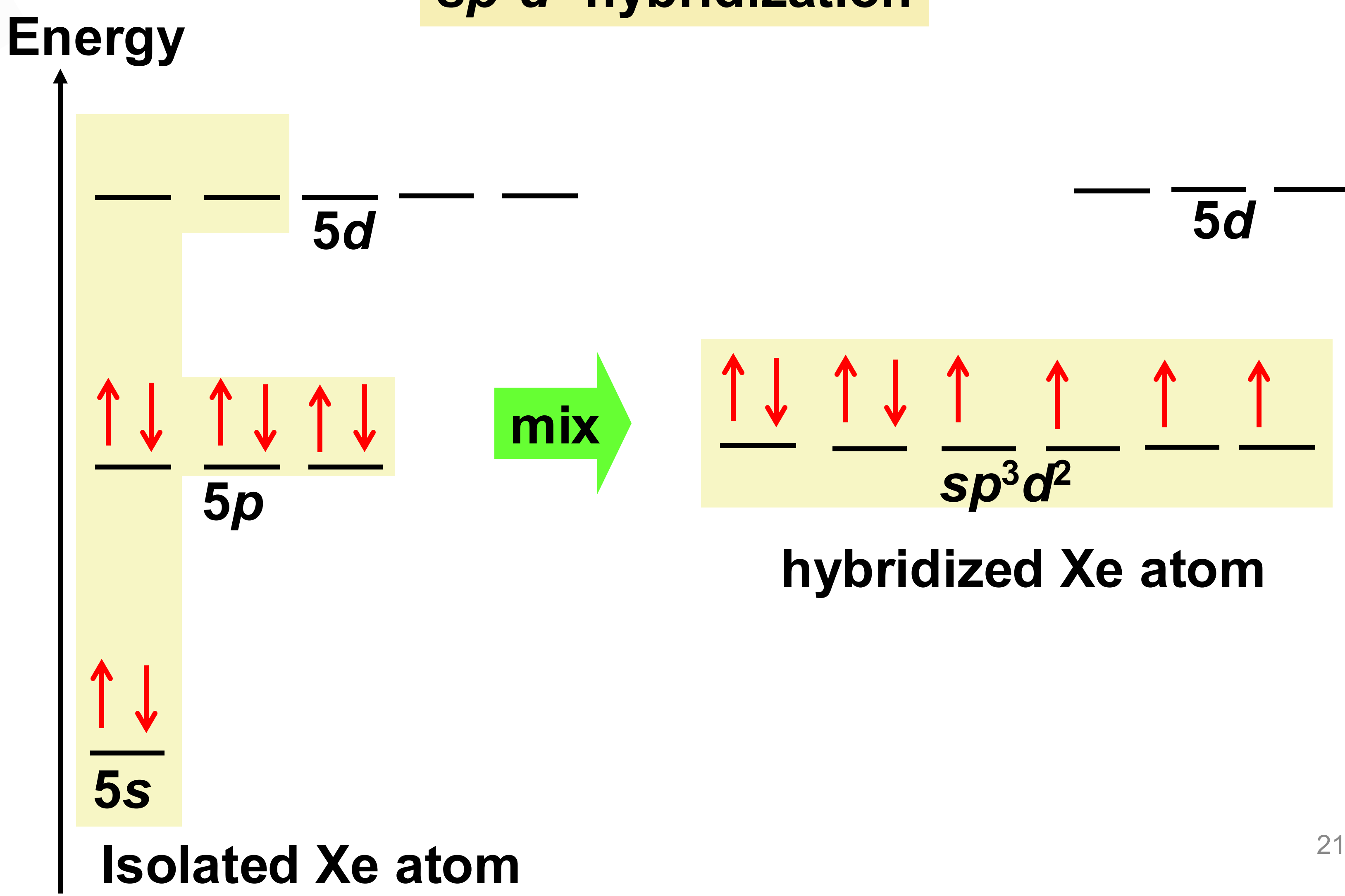
- Promotion of electron: ↑↓ (5s) | ↑↓ ↑ ↑ (5p) | ↑ ↑ _ _ _ (5d)

- sp^3d^2 hybridization: ↑↓ ↑↓ ↑ ↑ ↑ ↑ (sp^3d^2) | _ _ _ (empty 5d)

Ans: Example 5

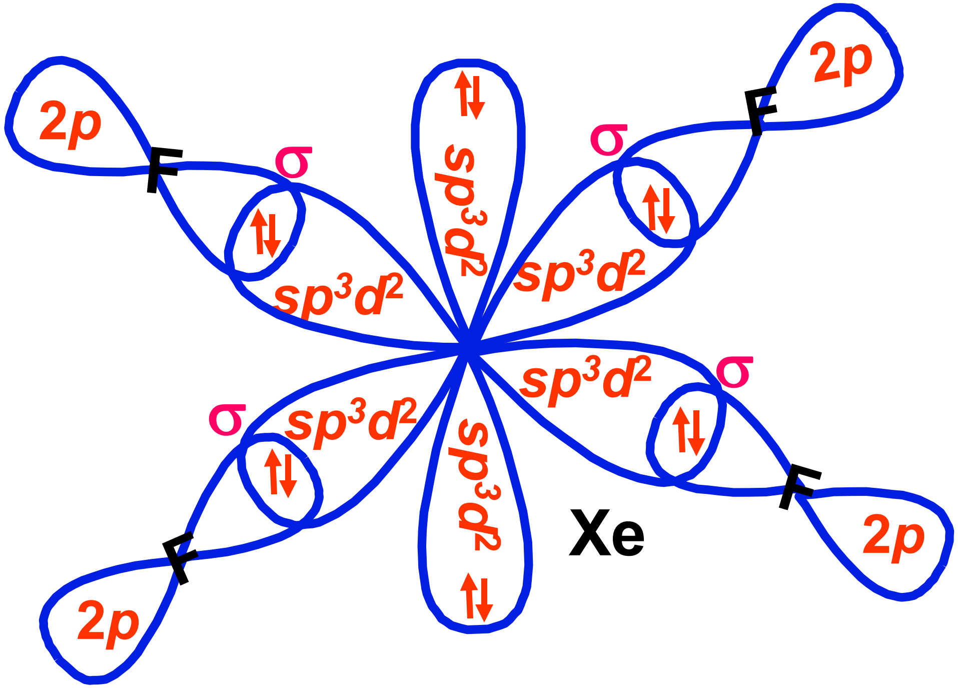
4.3

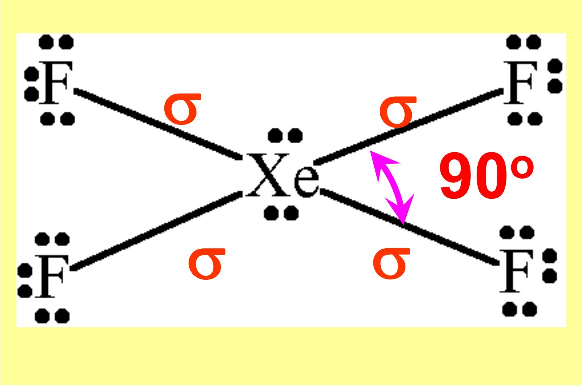
sp^3d^2 hybridization



Ans: Example 5

4.3





- **Electron group arrangement: octahedral**
- **Molecular shape: square planar**

Example 6

4.3

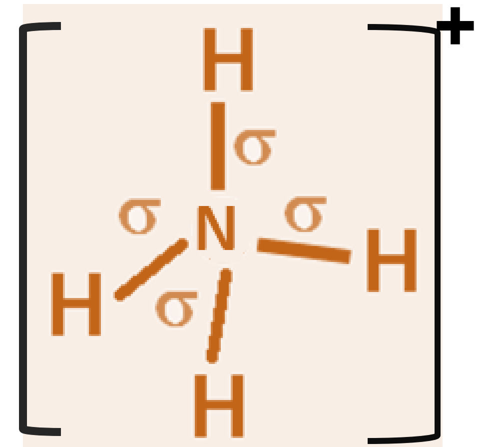
Describe a hybridization scheme for the central N atom in the molecule NH_4^+ that is consistent with the molecular shape.

Ans: Example 6

e^- configuration of N : $1s^2\ 2s^2\ 2p^3$

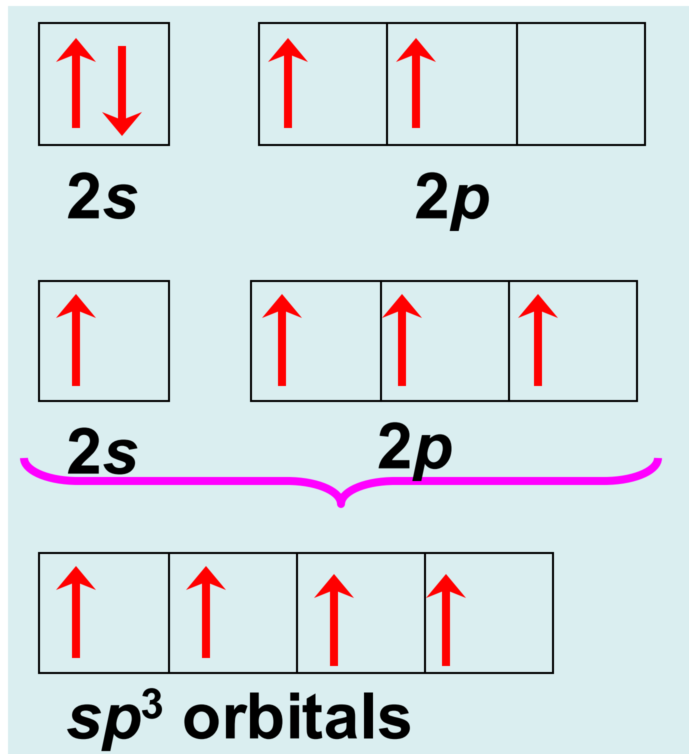
e^- configuration of N^+ : $1s^2\ 2s^2\ 2p^2$

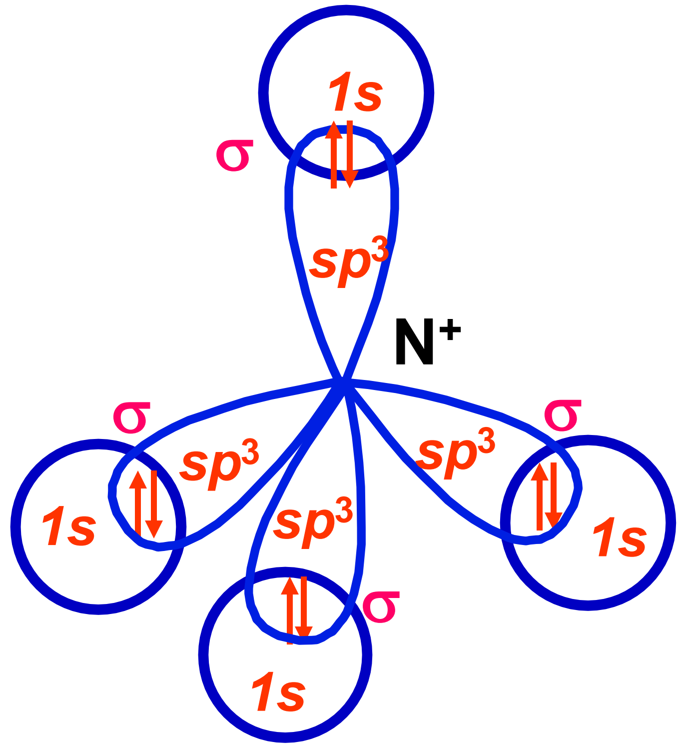
e^- configuration of H : $1s^1$

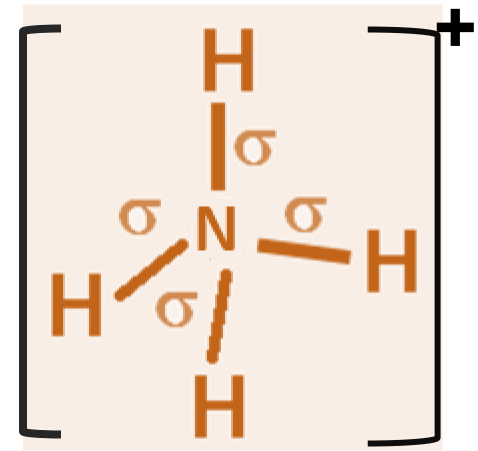


Valence e^- in N^+:

- **Ground state:**
- **Promotion of e^-:**
- ***sp*^{3} hybridization:**







- **Each N–H (σ) bond formed by: Overlap of one *sp*^{3} hybrid of N$^+$ and one 1*s* orbital of H atom**

- Four equivalent *sp*^{3} hybrid orbitals that lie 109.5^o apart
  ☞ four e$^-$ groups
  (from VSEPR theory)

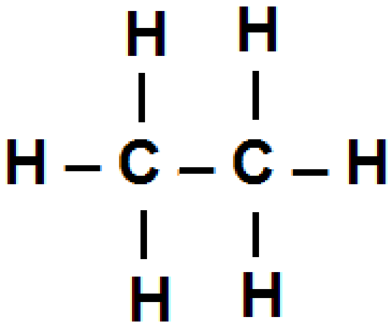
- e$^-$ group arrangement = tetrahedral

  Molecular shape = tetrahedral

MOLECULES CONTAINING MULTIPLE BONDS

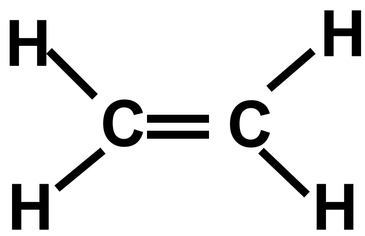
C_2H_6
ethane

- Tetrahedral
- H-C-H about 109.5°
- *sp*^{3}



C_2H_4
ethylene

- Trigonal planar
- H-C-H is 120°
- *sp*^{2}



C_2H_2
acetylene

- linear
- H-C-H is 180°
- *sp*

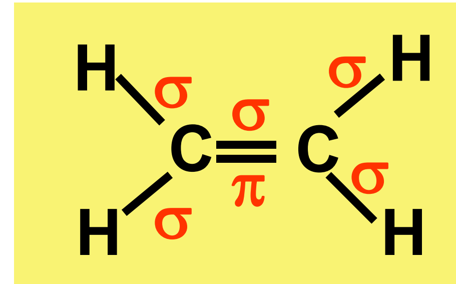
$H—C \equiv C—H$

BONDS IN C_2H_4

Ethylene ($CH_2{=}CH_2$)

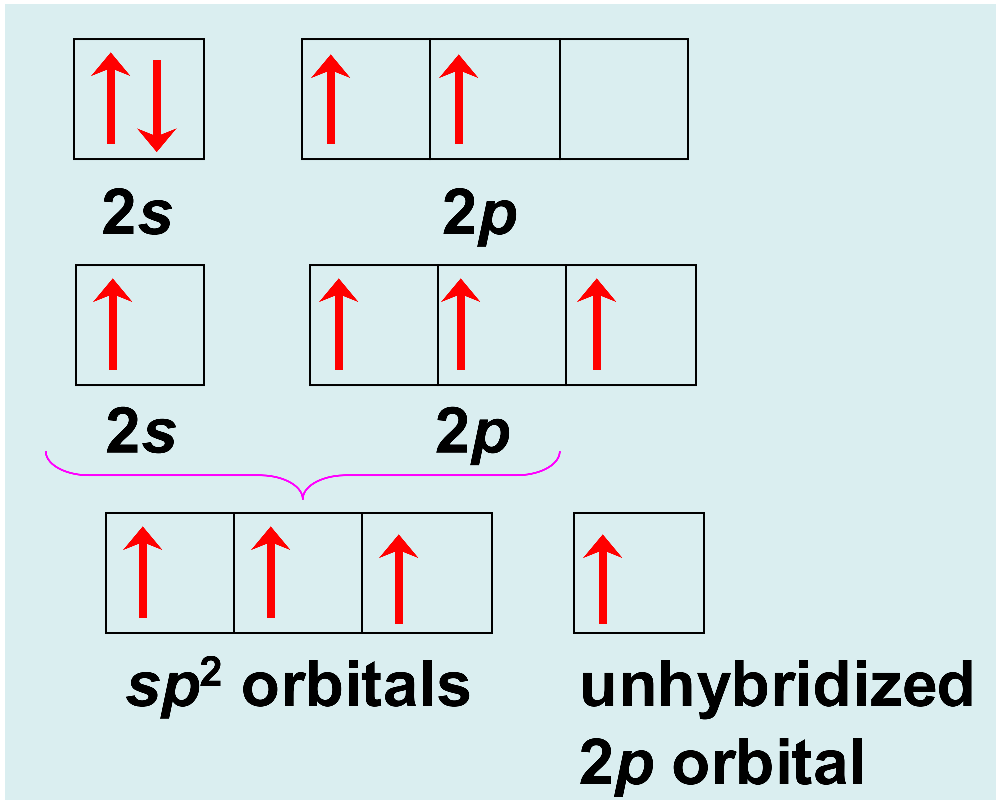
e^- configuration of C : $1s^2\ 2s^2\ 2p^2$

e^- configuration of H : $1s^1$

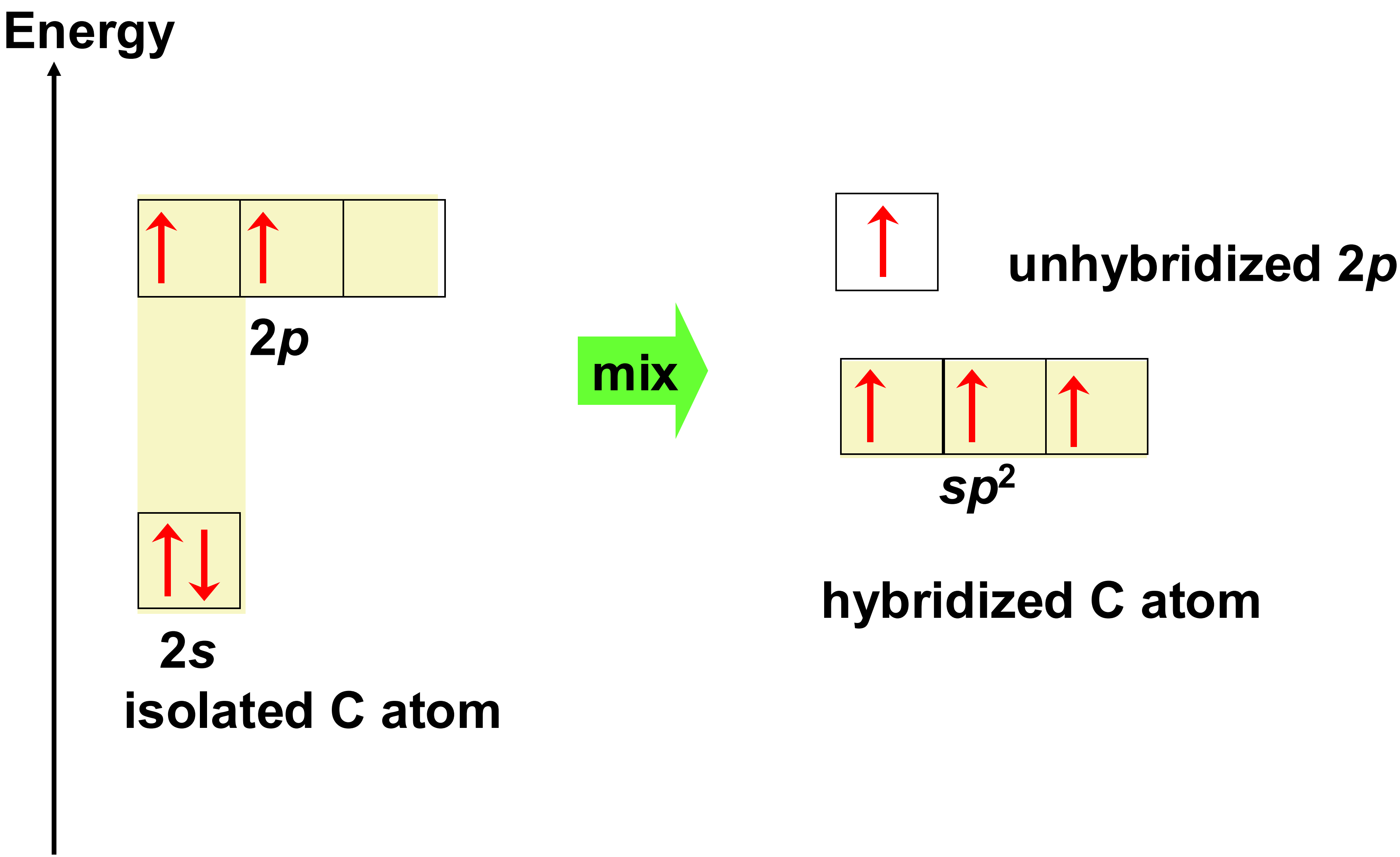


Valence e^- in C:

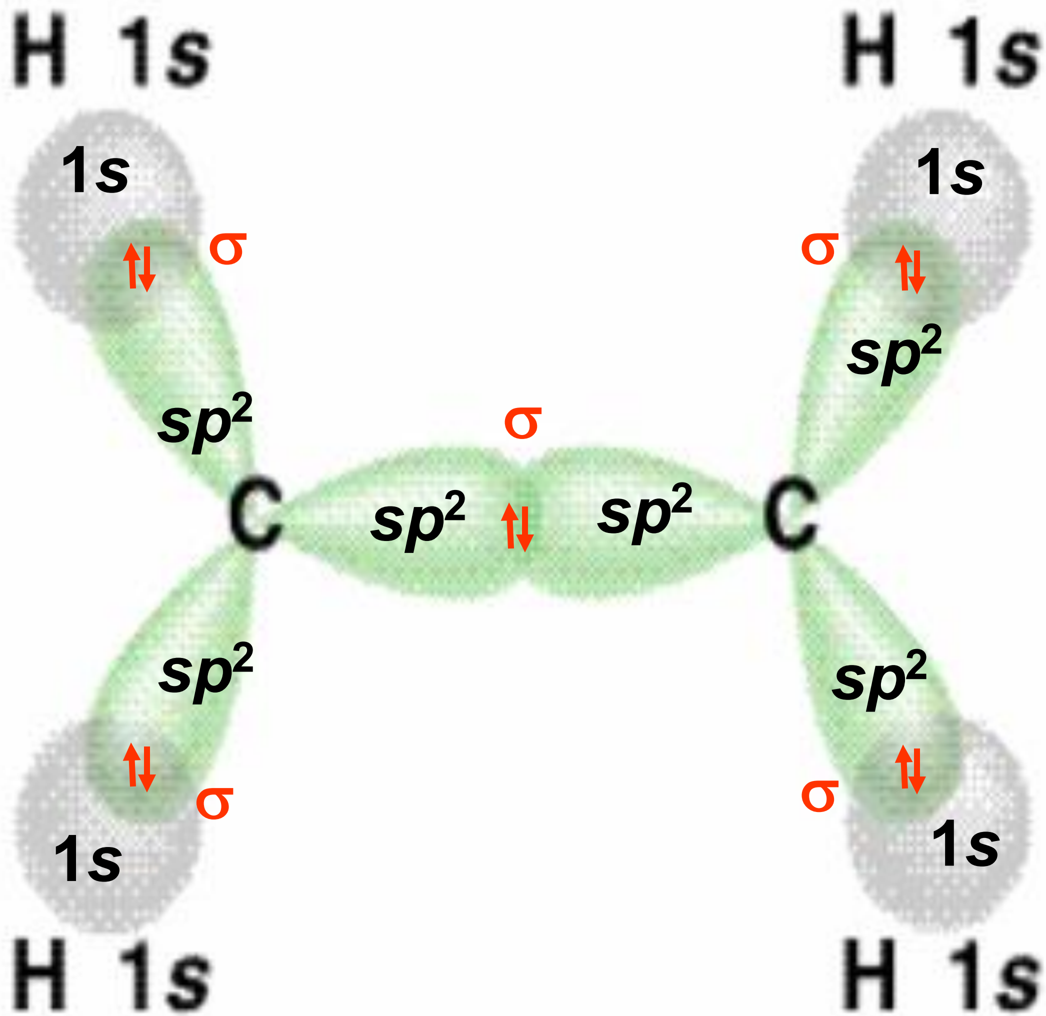
- Ground state:
- Promotion of e^-:
- *sp*^{2} hybridization:

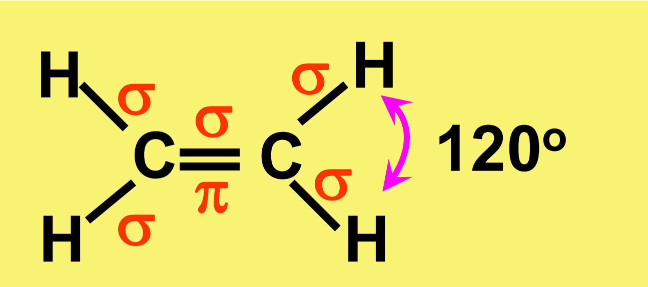


sp^2 hybridization

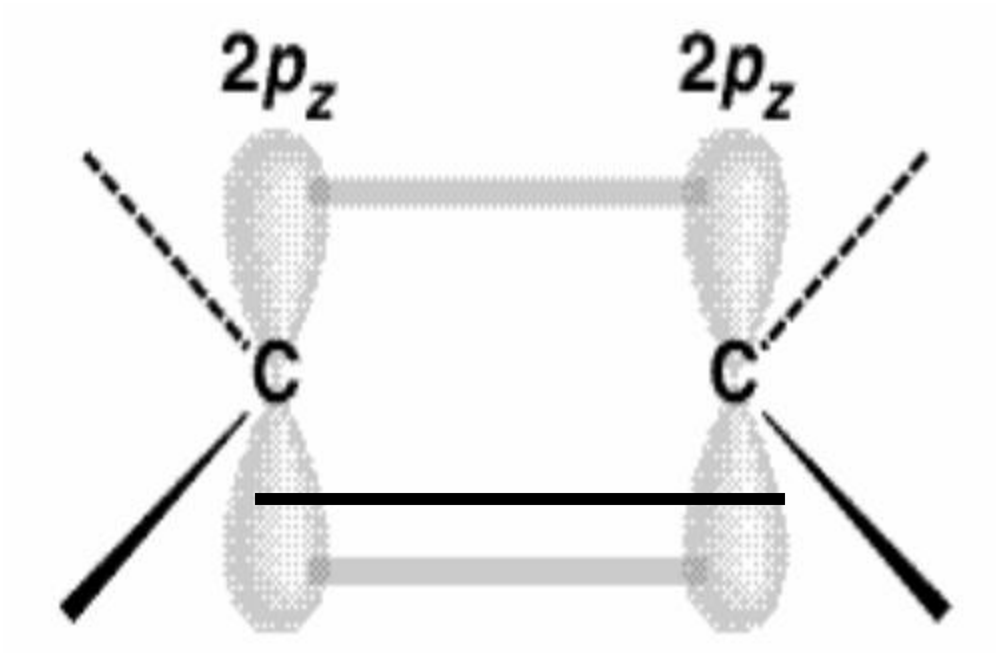


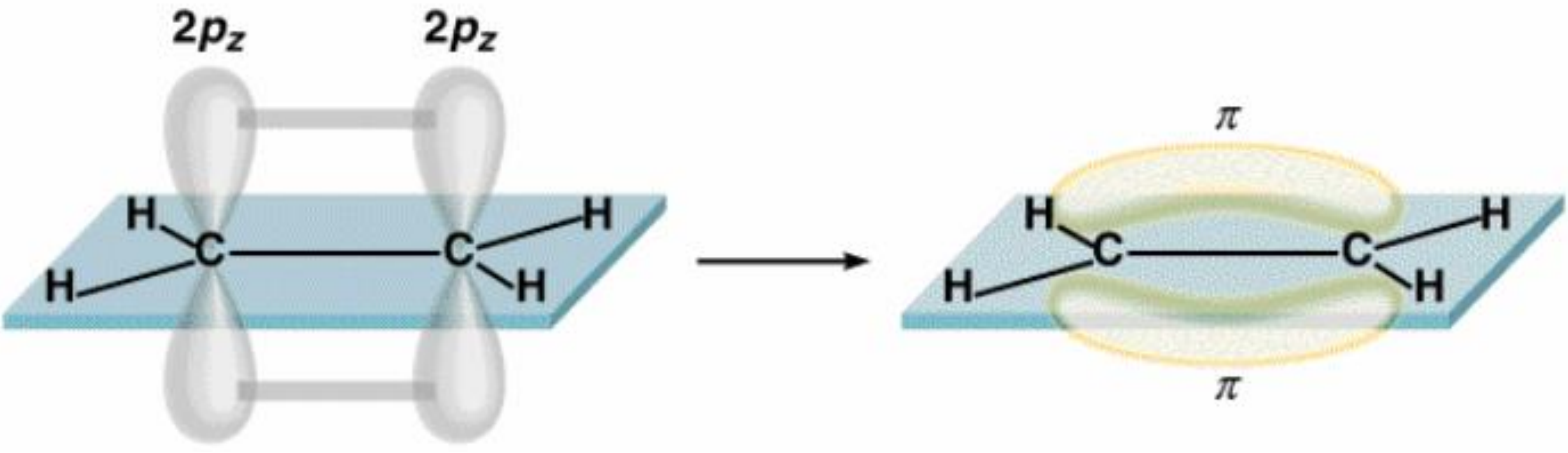
σ Framework in Ethene (C_2H_4)





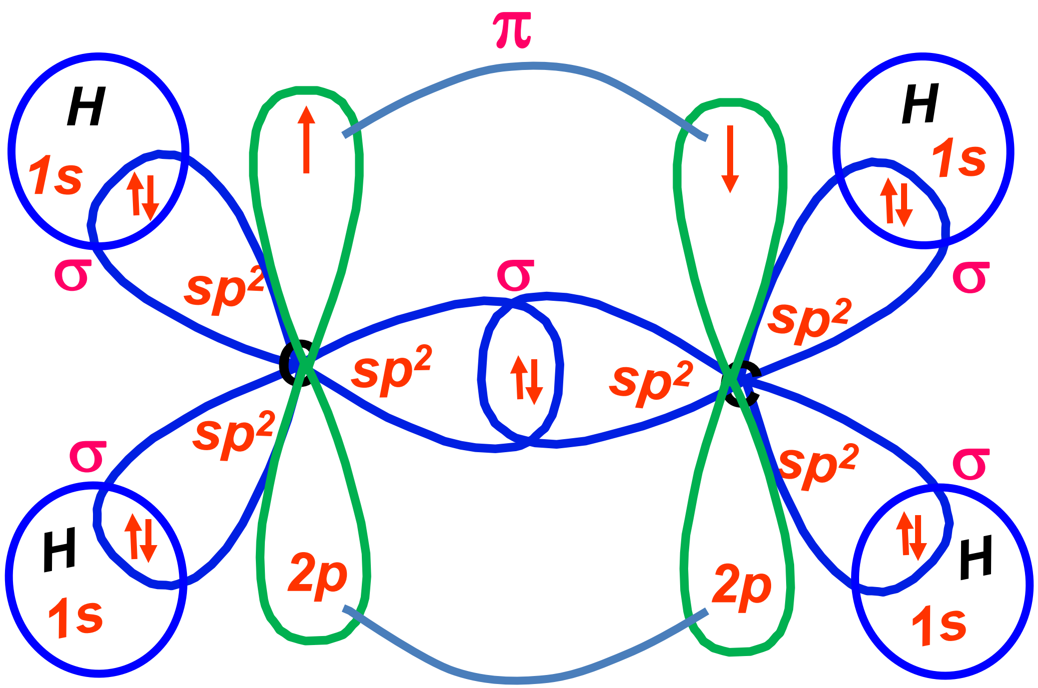
π Bond in Ethene (C_2H_4)

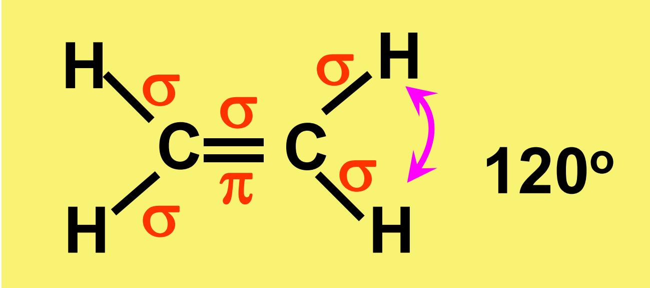




- All 6 atoms are in the same plane

- Overlap of the 2*p* orbitals restricts rotation of C–C bond and cause ethylene takes planar structure





- **σ–bond**:
  C–H: overlap *sp*² of C with 1*s* of H
  C–C: overlap of *sp*² of both C

- **π-bond in C=C bond:**
  Overlap of unhybridized 2*p* of C with another unhybridized 2*p* of C

CHAPTER 4.4 : OVERVIEW

- **Describe IMF**
  - van der Waals
    - Dipole-dipole forces
    - London forces
  - Hydrogen bonding
- **Factors that influence strength**
  - van der Waals
    - Dipole moment (DD)
    - Polarizability/size (LF)
    - Molecular shape (LF)
  - Hydrogen bond
    - No. of HB
    - Electronegativity
- **Effects of HB**
  - Boiling point
  - Solubility
  - Density of water

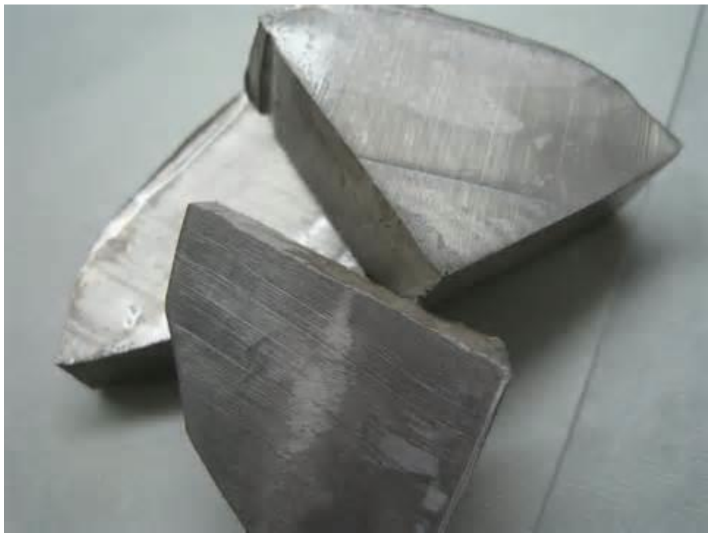
4.4 INTERMOLECULAR FORCES

Teaching and learning outcomes

At the end of the lesson, student should be able to

4.4	**Intermolecular forces**
a)	Describe intermolecular forces (C1, C2) i. van der Waals forces • Dipole-dipole interactions or permanent dipole • London forces or dispersion forces ii. Hydrogen bonding
b)	Explain factors that influence van der Waals and hydrogen bond. (C2, C3)
c)	Relate the effects of hydrogen bonding on the following physical properties: (C2, C3) i. Boiling point ii. Solubility iii. Density of water compared to ice

How does this substance melt ?



heated

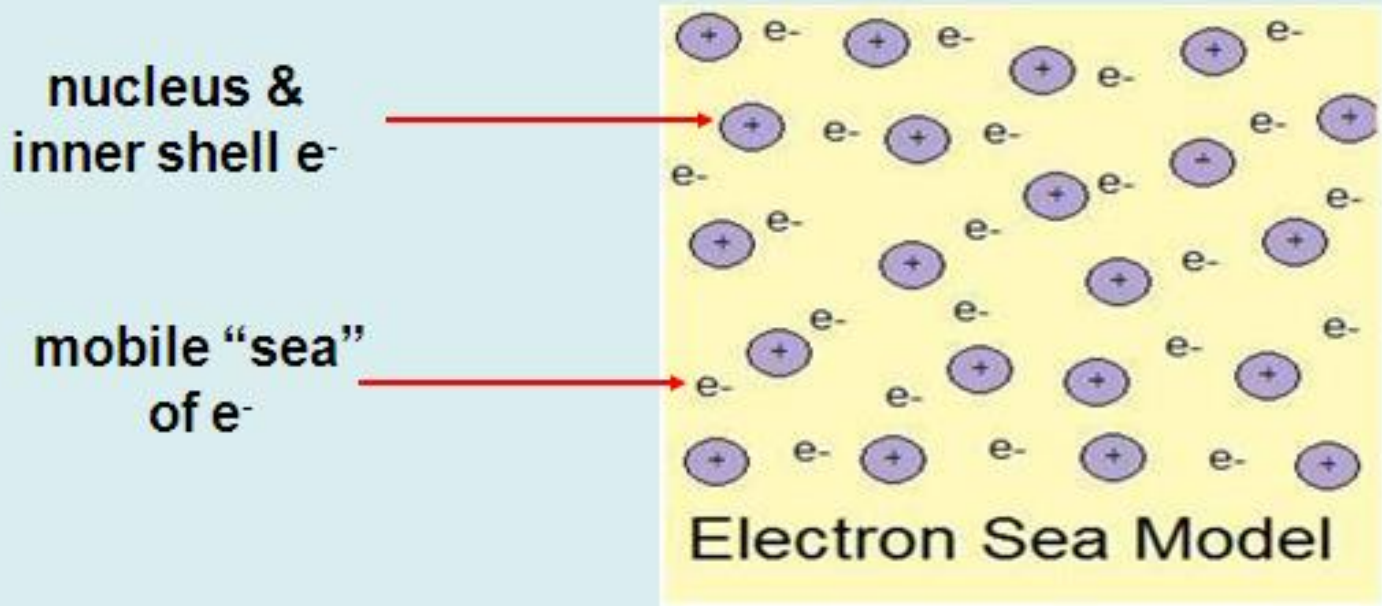
metal

Metal
melting

METAL

- Metals have metallic structure where the atoms are closely packed with each other
- Metallic bond is formed from the electrostatic attraction between the positively charged metal ions and the sea of delocalised valence electrons
- Metals also have very high melting and boiling point because they have strong metallic bonding

Cross Section of a Metallic Crystal



Metal	Melting point (ºC)
Na	98
Mg	650
Al	660

IONIC COMPOUND

- Made up **positive and negative ions**

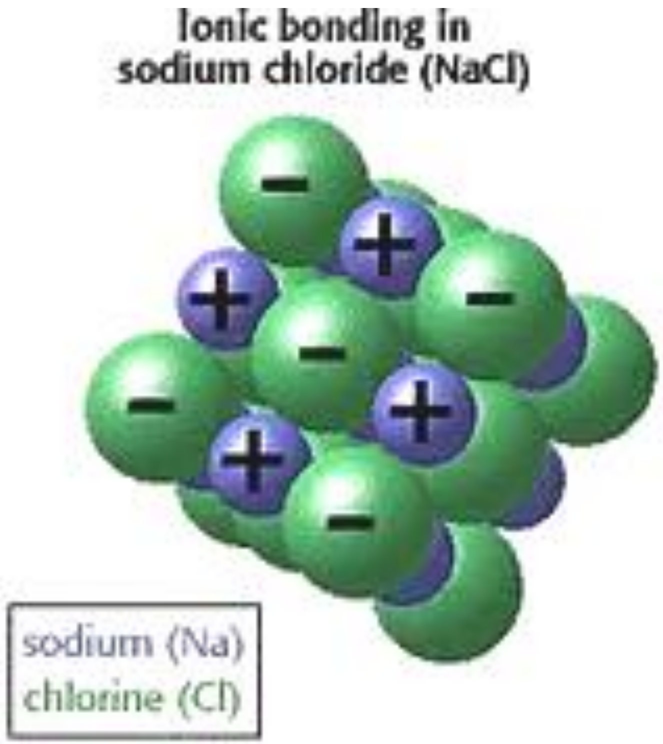
- The oppositely charged ions are formed through the **transferring of electrons from metal to non-metal**

- **Ionic bond** is formed from the electrostatic attraction between positive ion and negative ions in an ionic compound.

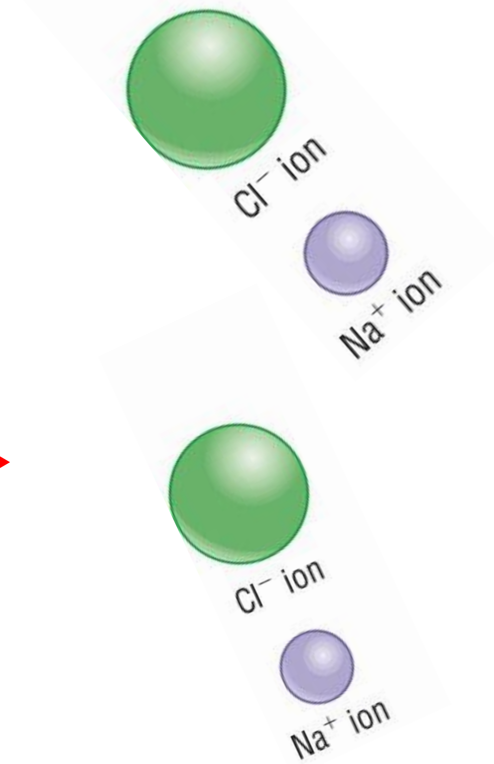
- **Examples : $NaCl$, KCl, $MgBr_2$**

- The **electrostatic attraction (ionic bond)** between **positive and negative ions** is **strong**. It takes a **lot of energy to overcome this attraction** for the ionic compounds to melt.
- Ionic compounds have **higher melting point**.

Ionic Compounds	Melting Point (ºC)
$NaCl$	801
CaF_2	1423
MgO	2852



melting



Solid NaCl

Liquid NaCl

COVALENT COMPOUND

- A covalent bond is formed when one or more nonmetals combine with each other to form molecules

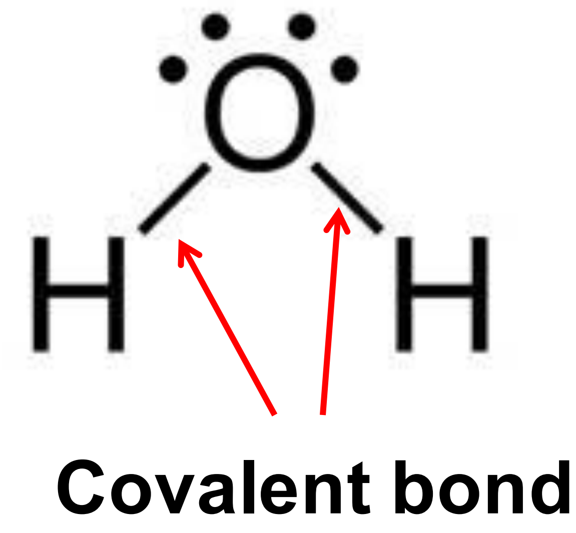
- Atoms (nonmetals) bonded by shared electron pairs

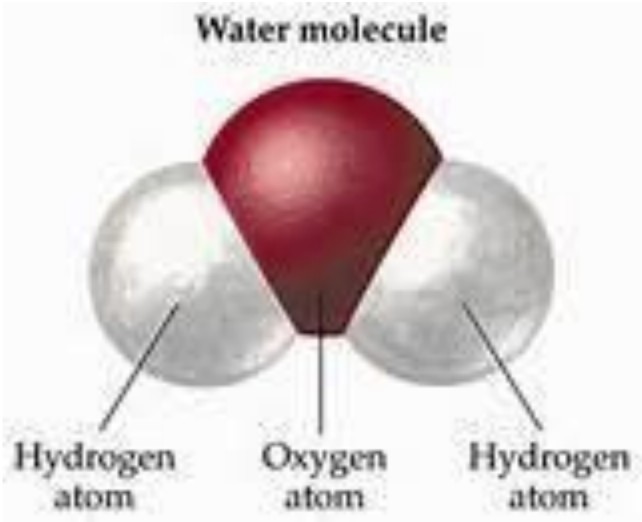
- Examples : H_2, O_2, CO_2 , CH_4, Si

Simple covalent molecular structure

Water, H_2O

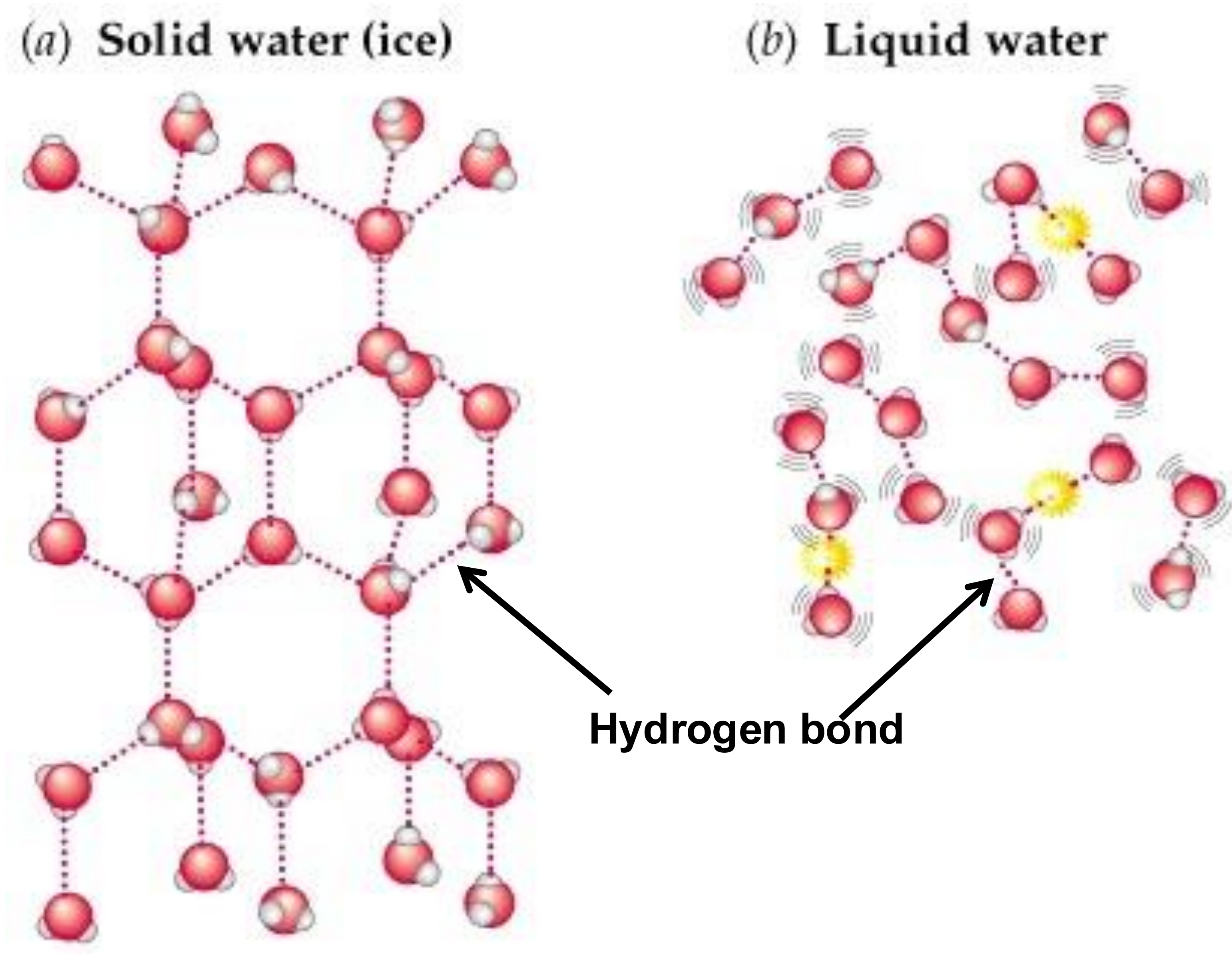
In water (H_2O), O and H atoms are held by strong covalent bonds

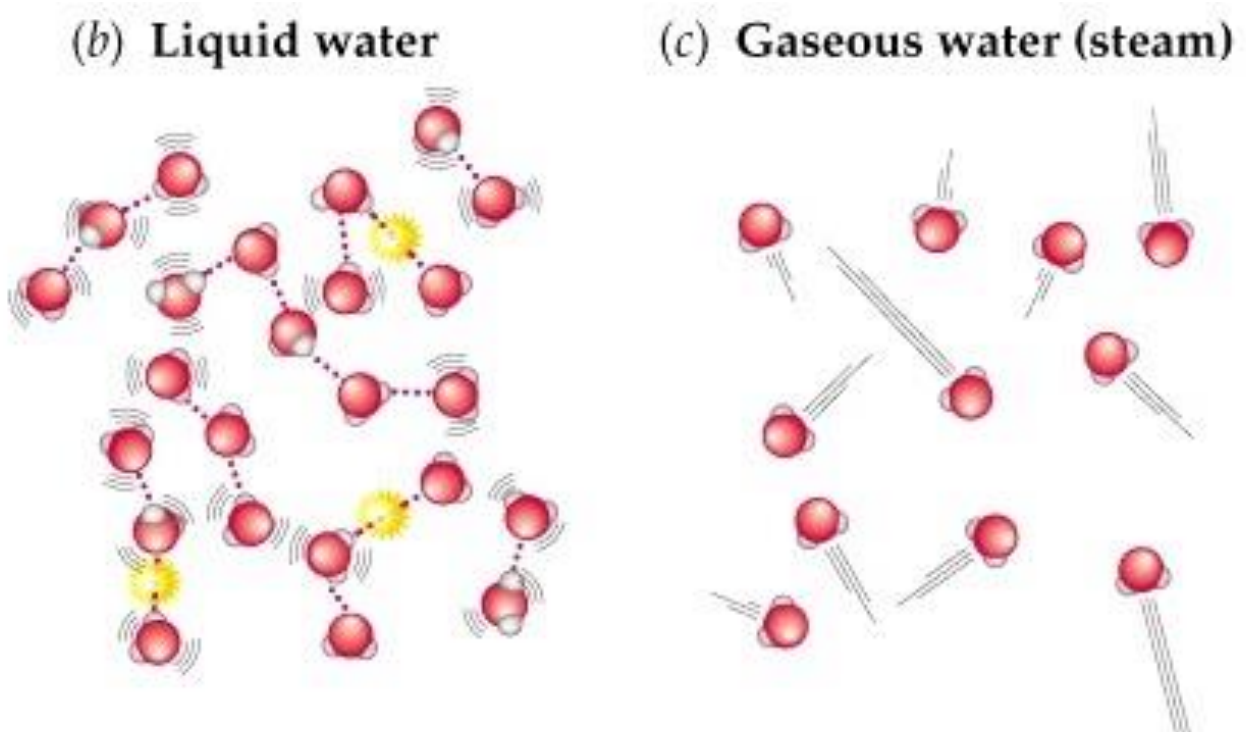




- The water molecules are held together by **stronger forces** which is called **hydrogen bond**

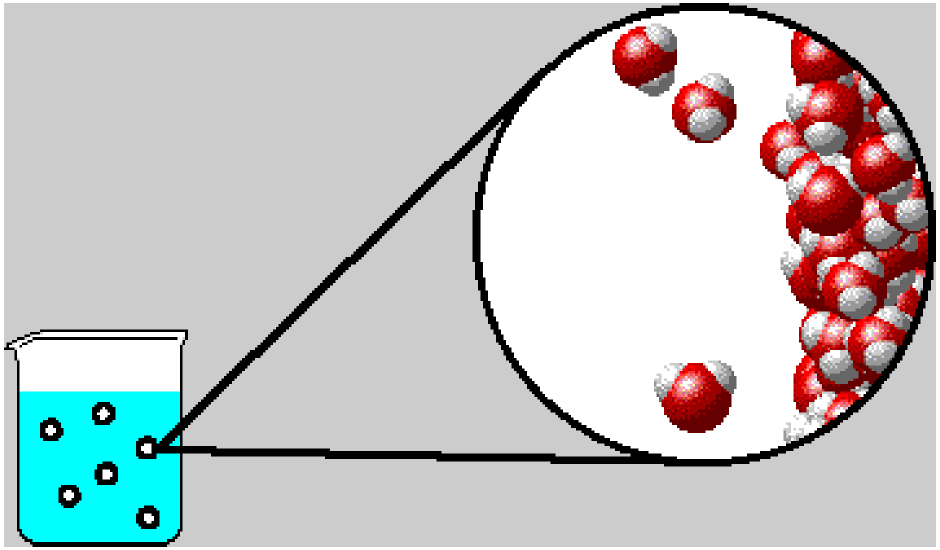
- During melting process, some of the hydrogen bonds are broken. The solid ice (water) changed to liquid water


(a) Solid water (ice)
(b) Liquid water
Hydrogen bond



- **During boiling process, water molecules (in liquid water) need to overcome the stronger hydrogen bonds exist between them and finally change to gaseous water (water vapour)**

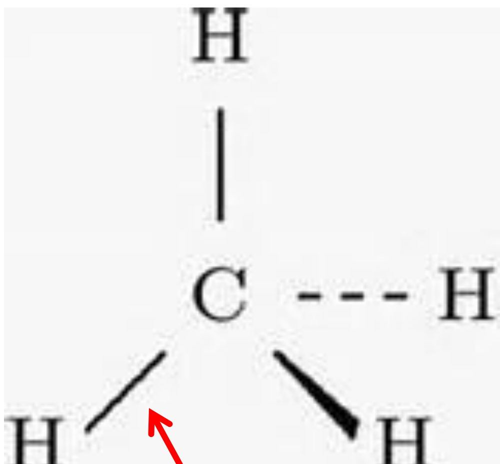
Molecular size (molar weight) = 18 g
Boiling point of water = 100°C



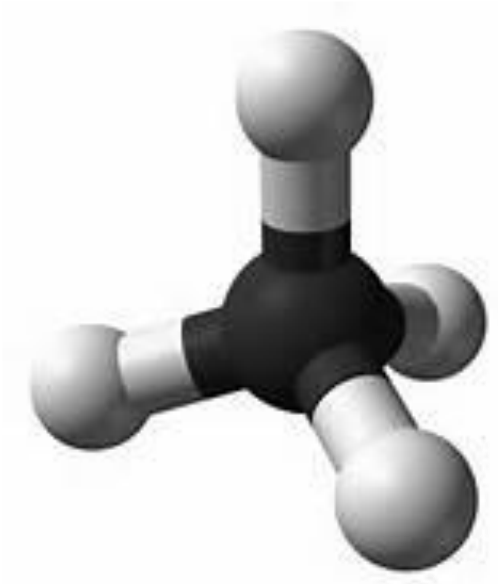
Simple covalent molecular structure

Methane, CH_4

In methane (CH_4), C and H atoms are held by strong covalent bonds



Covalent bond



- **CH_4 molecules are held together by weak van der Waals forces between its molecules.**

- **During boiling process, CH_4 molecules need to overcome the forces**

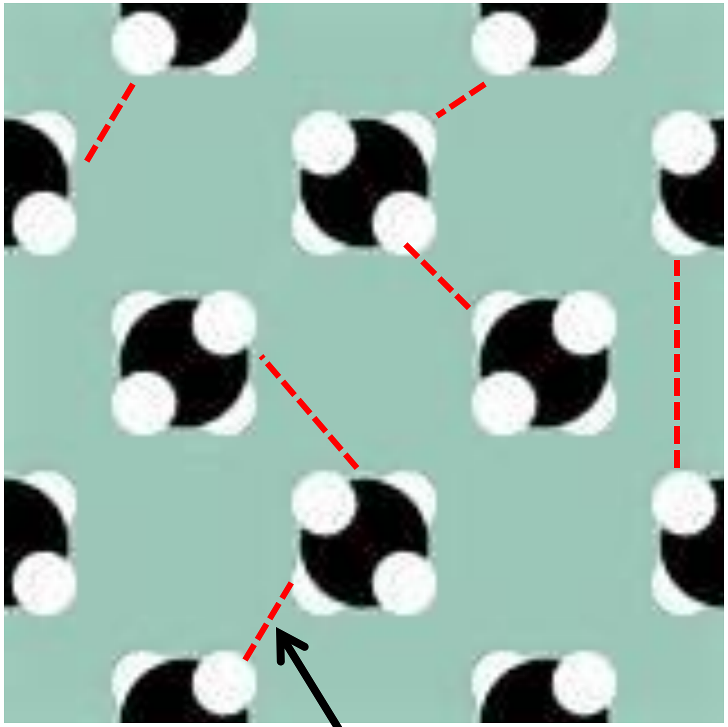
- **Less energy needed to overcome the weak van der Waals forces.**

- **Therefore, CH_4 has lower boiling point**

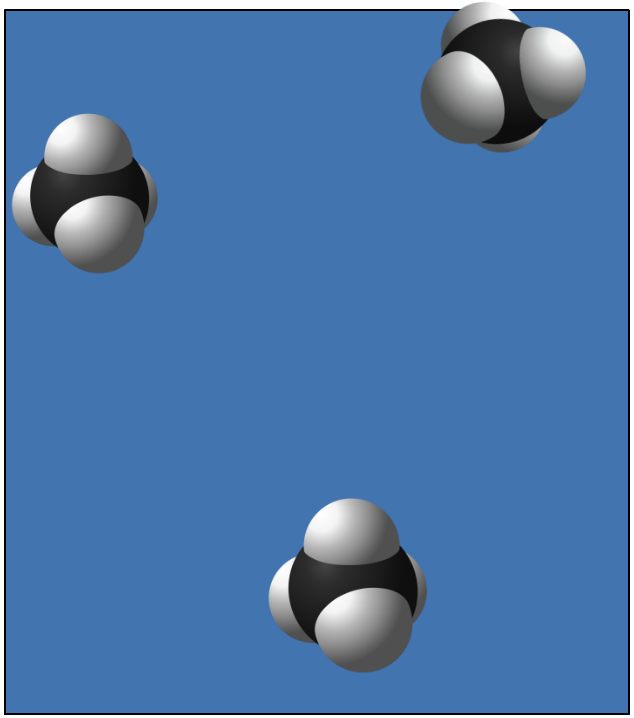
Simple covalent molecular structure

Methane, CH_4

Liquid CH_4

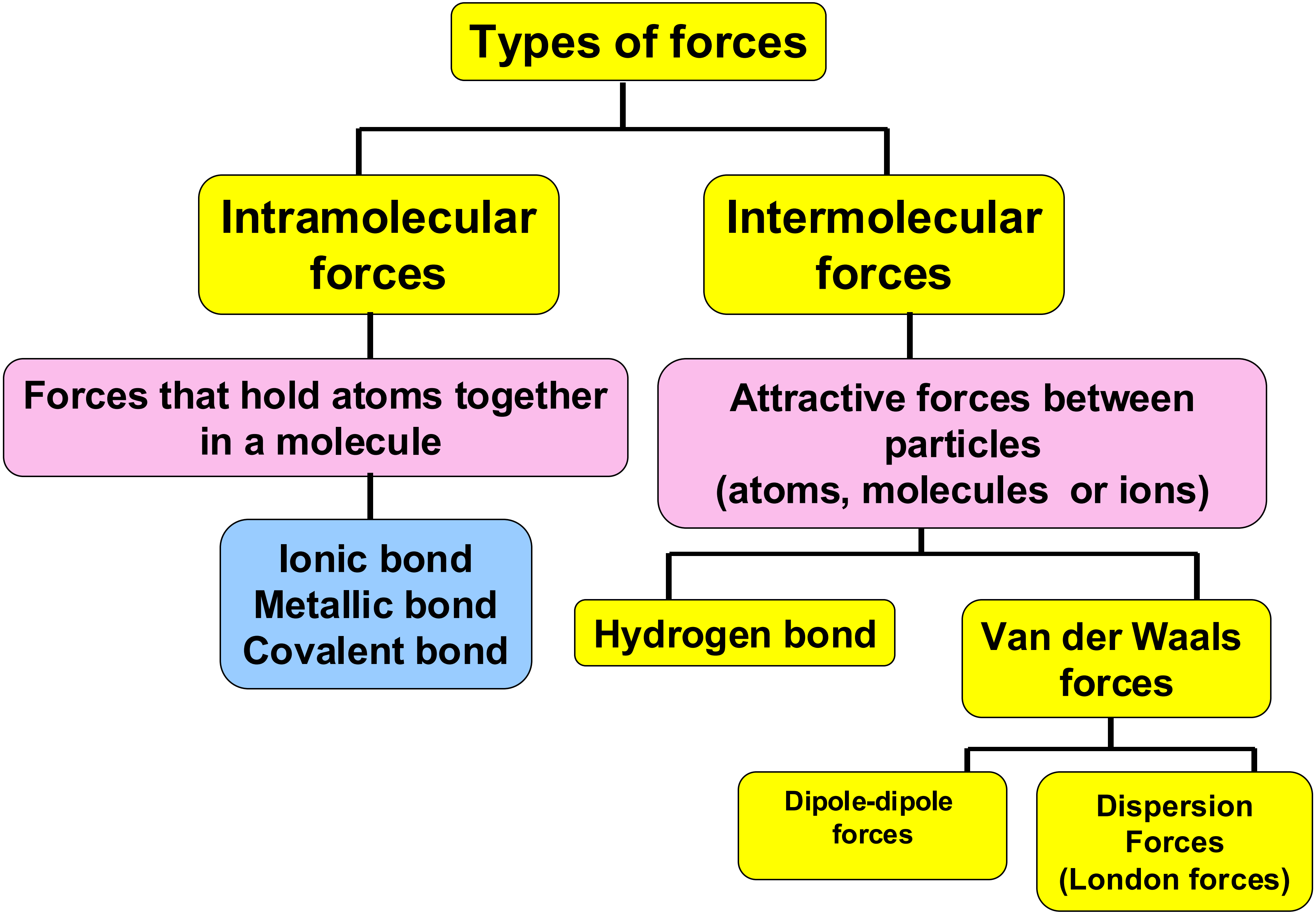


Gaseous CH_4



Van der Waals forces

Molecular size (molar weight) = 16 g
Boiling point of CH_4 = -161.49°C


Types of forces
Intramolecular forces
Intermolecular forces
Forces that hold atoms together in a molecule
Attractive forces between particles (atoms, molecules or ions)
Ionic bond
Metallic bond
Covalent bond
Hydrogen bond
Van der Waals forces
Dipole-dipole forces
Dispersion Forces (London forces)

INTERMOLECULAR FORCES

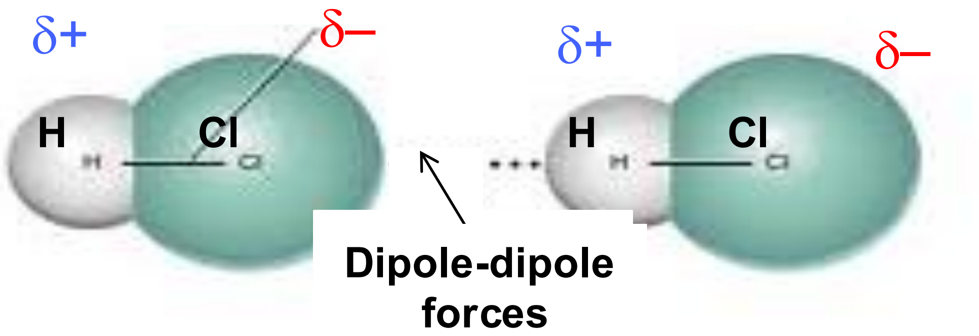
Attractive forces between neighboring particles (atoms, molecules or ions).

Exist when the molecules are sufficiently close to each other.

❶ van der Waals Force

(a) DIPOLE–DIPOLE (DD) FORCES

- ❑ Attractive forces between polar molecules
- ❑ The positive pole of one molecule attracts the negative pole of another



STRENGTH OF DIPOLE-DIPOLE FORCES

For a compound of similar molecular size:

- **the more polar the molecule, the stronger the strength of its dipole-dipole forces**

- **more energy is needed to overcome the forces between molecules**

❶ van der Waals Force

(b) LONDON FORCES (DISPERSION FORCES)

- **Dispersion forces are caused by the motion of electron in a molecule or atom**

- **The forces are weak, but exist in any particles (non polar, polar molecule, ion, atom,etc.)**

HOW DO THE LONDON FORCES ARISE BETWEEN MOLECULES OR ATOMS ?

① **The electrons are distributed uniformly around the nucleus**

② **At any instant, there may be more electrons on one side of the nucleus**
☞ **instantaneous dipole**

③ **When the atoms are closed together, the instantaneous dipole in one atom induces a dipole in its neighbors**
☞ **induced dipole**

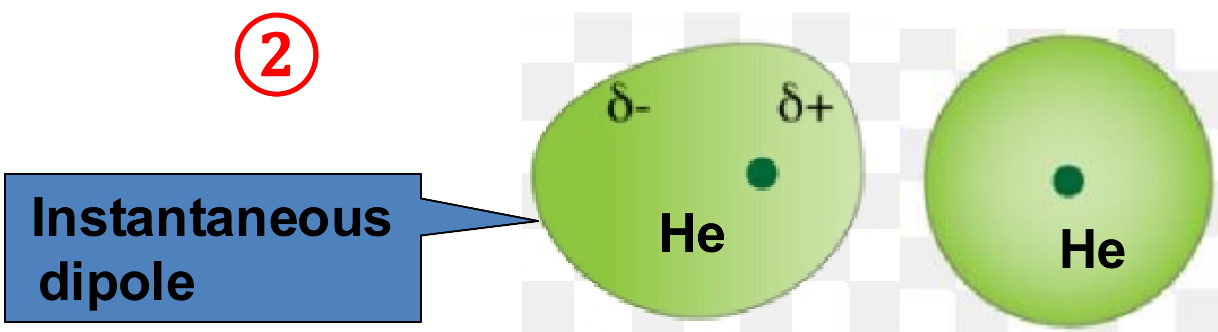
④ **Attraction between instantaneous dipole–induced dipole forms London Force**

Example :
London Forces (Dispersion forces) between He atoms

①

A (nonpolar) **B (nonpolar)**

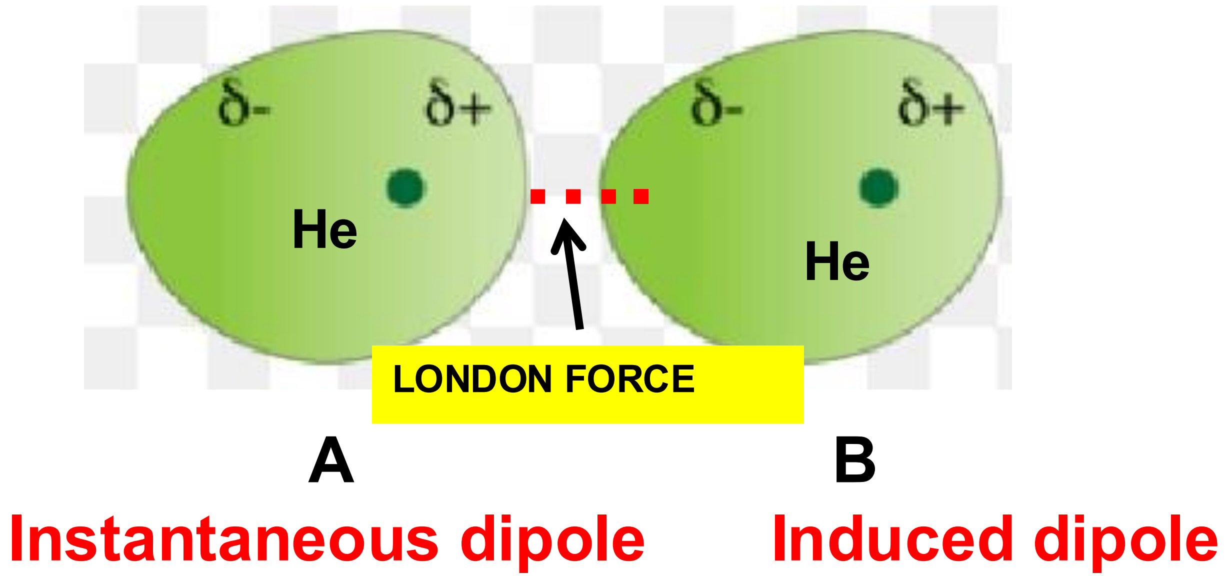
②



A **B (nonpolar)**

Instantaneous dipole on atom A is caused by uneven distribution of electrons

③

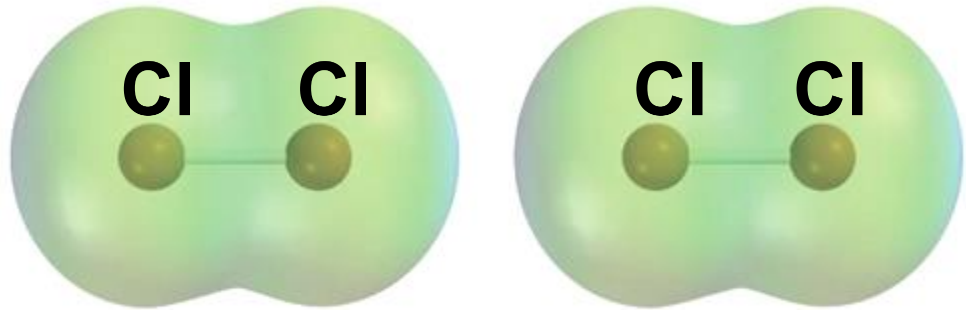


The instantaneous dipole on atom A induces the neighboring atom when they are very close.

The London force occurs as neighboring atoms attract one another. It is significant when atoms or molecules are very close to each other.

Example :
London Force (Dispersion forces) between Cl_2 molecules

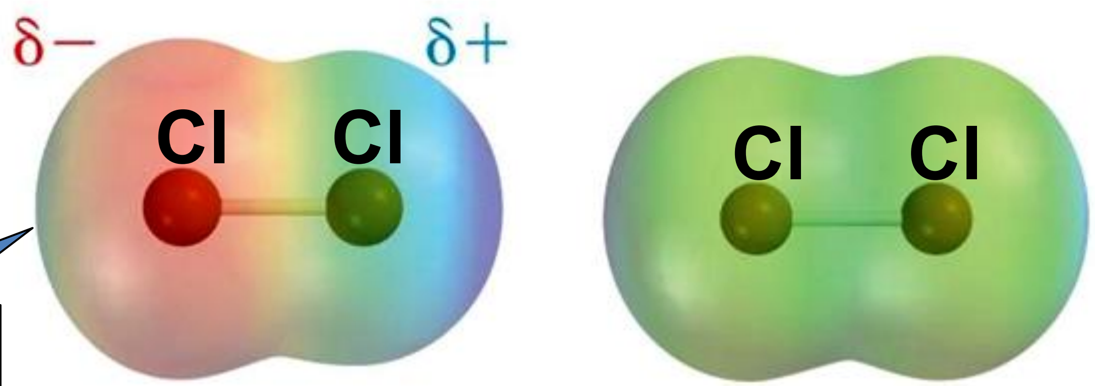
①



Molecule A (nonpolar)

Molecule B (nonpolar)

②

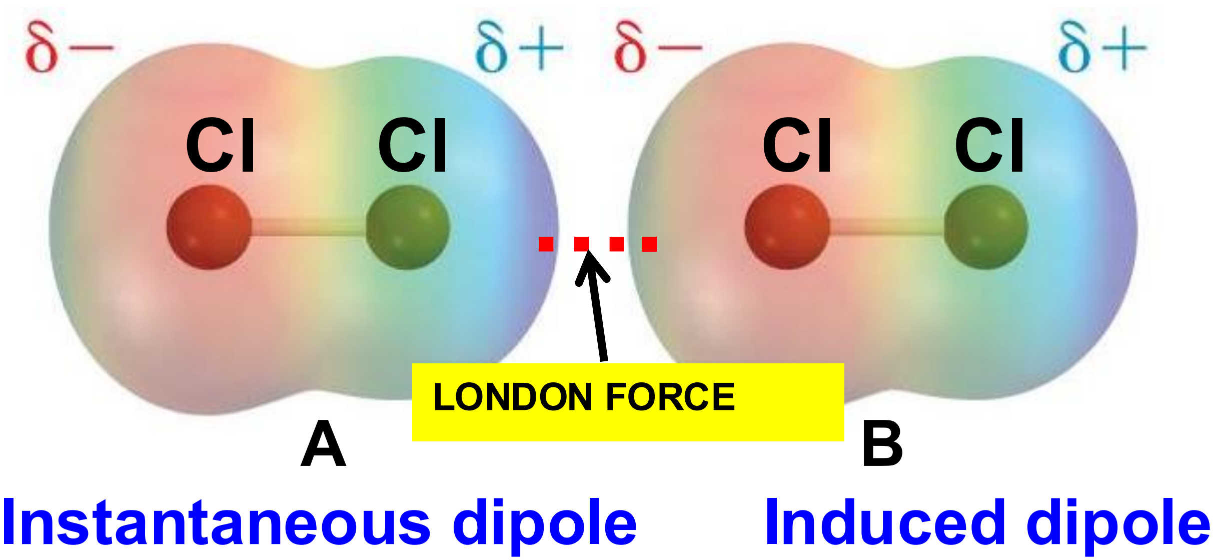


Instantaneous dipole

A

B (nonpolar)

At any instant, the molecule has an instantaneous dipole

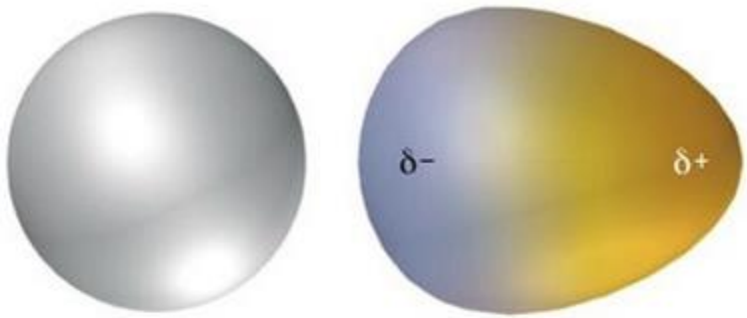


Instantaneous dipole on molecule A induces the neighbouring molecule when they are very close, causing the molecule to be attracted to each other.

This attractive interaction is called London force

STRENGTH OF LONDON FORCE

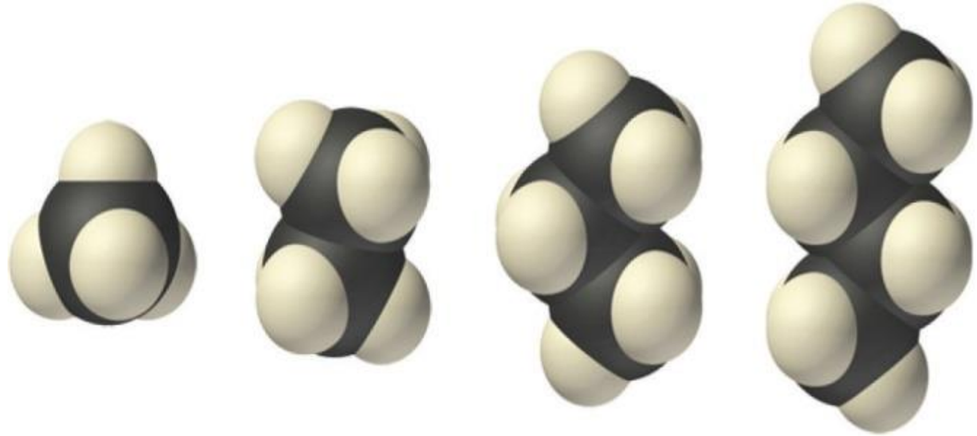
- Influenced by:
  - Polarizability (size)
  - Molecular shape



❖ Polarizability (Size)

- Depends on number of e^-, which correlates closely with size / molar mass

  - Molecules with bigger size / larger molar mass,
  - Has more number of electrons,
  - Polarizability of molecule increases,
  - Stronger London force



Methane	Ethane	Propane	*n*-Butane
16 g/mol	30 g/mol	44 g/mol	58 g/mol
−161.5°C	−88.6°C	−42.1°C	−0.5°C

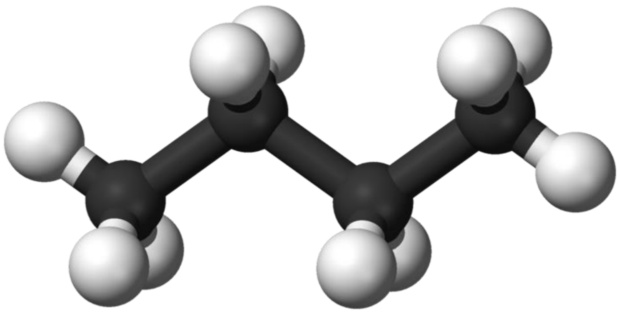
Effect of polarizability (size) on boiling point of molecules

- **Methane, ethane, propane and butane are nonpolar molecules**
- **London forces exist between their molecules respectively**
- **The strength of London forces is directly proportional to the polarizability / (molecular size).**
- **More energy needed to overcome the London forces between butane molecules. Thus, butane has the highest boiling point.**

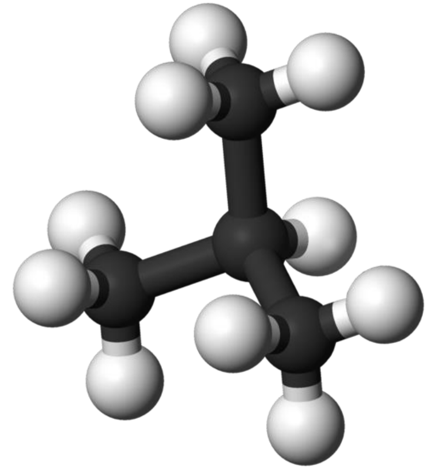
❖ Molecular Shape

- **For nonpolar substance with the same molar mass:**

  - **larger surface area of molecules**
  - **increase contact between molecules**
  - **Stronger London force**



C_4H_{10}
Butane
(molecular weight = 58 g)
Boiling point = −0.5 ºC



C_4H_{10}
2−methylpropane
(molecular weight = 58 g)
Boiling point = −11.7 ºC

Effect of molecular shape on boiling point of molecules

- **Butane and 2–methylpropane are nonpolar molecules and have same molecular size**
- **The surface area of butane is larger than 2–methylpropane. Thus, it can form more London forces between molecules.**
- **The London forces between butane molecules is stronger**
- **More energy required to overcome the stronger forces. Thus, butane has the highest boiling point**

London Force VS Dipole-Dipole Force

- When the molar mass (molecular size) of a nonpolar molecule is larger than the polar molecule, London forces will be stronger than the dipole-dipole forces.

EXAMPLE:

CHF_3	CCl_4
M_w : 70 g	M_w : 154 g
polar	non polar
LF and DD	LF

bp = – 78.4°C

bp = 76.5°C

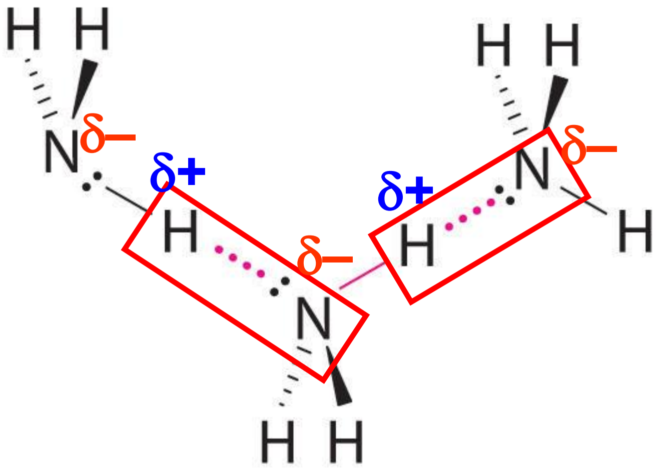
❷ HYDROGEN BOND

- **Force between polar molecules which have :**

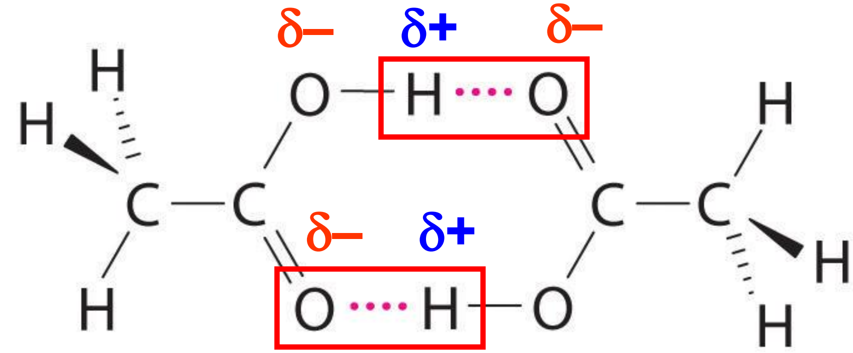
    1. **Partially positive H atom bonded to highly electronegative atom N, O or F**

        N–H, O–H or F–H

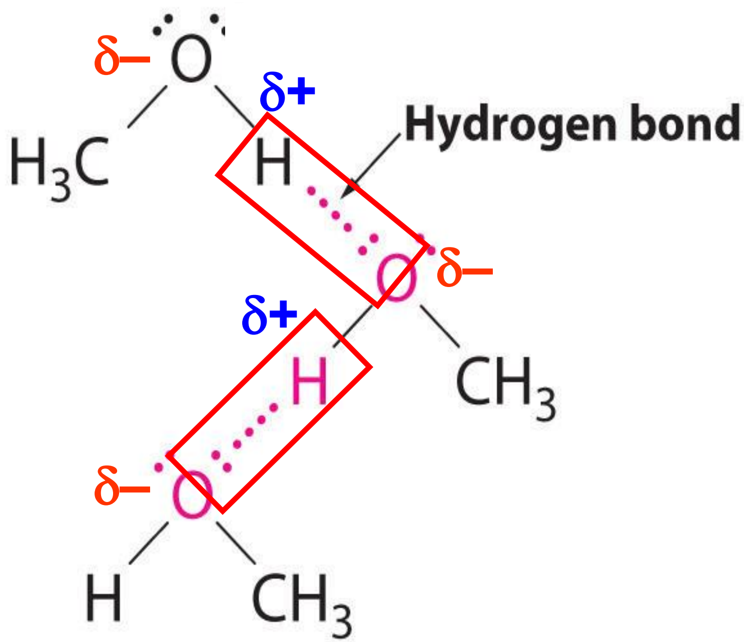
    2. **Partially negative lone pair electrons on N, O or F of another molecules**



Hydrogen bonding in ammonia



Hydrogen bonding in acetic acid



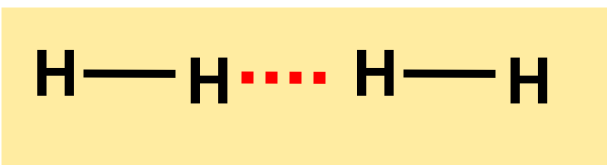
Hydrogen bonding in methanol

Example 1

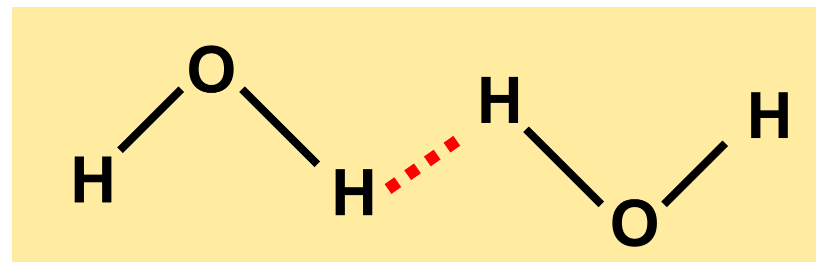
4.4

Which is the correct example of hydrogen bonding?

(a)

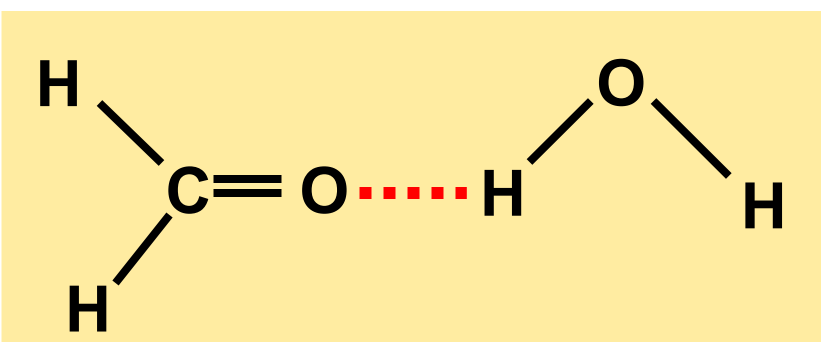


(b)

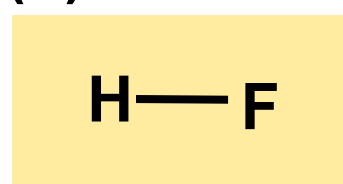


(c)

correct



(d)



Example 2

4.4

Show the formation of Hydrogen bonds between the each of the molecules given below.

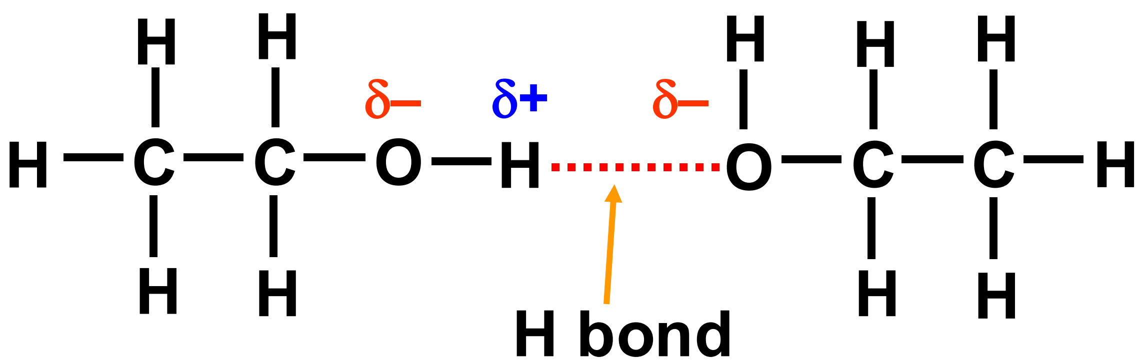
(a) CH_3CH_2OH

(b) $CH_3\underset{\underset{\displaystyle O}{\|}}{C}{-}NH_2$

Ans: Example 2

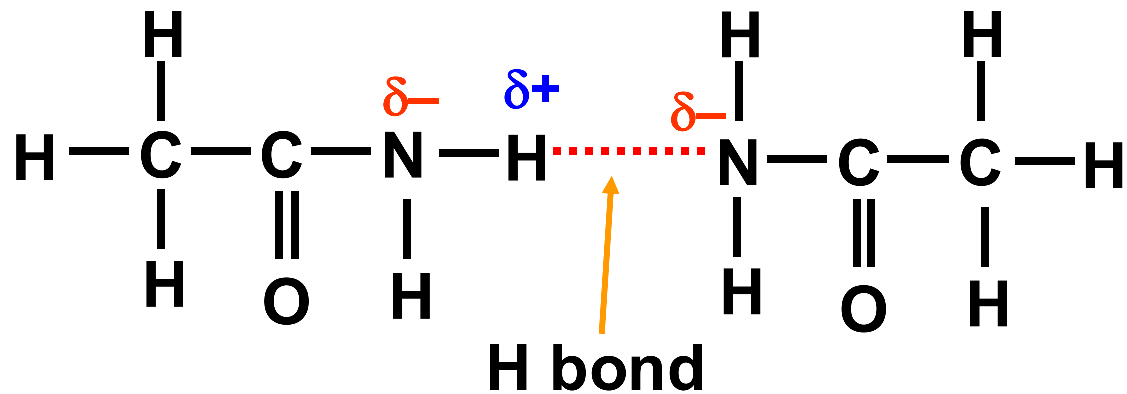
4.4

(a) CH_3CH_2OH

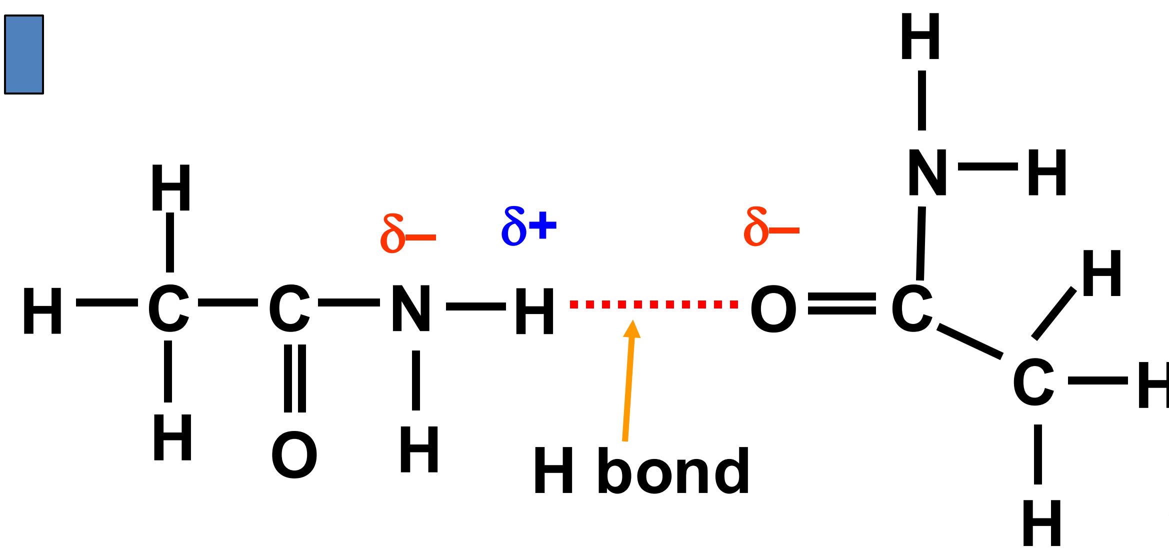


Ans: Example 2

(b) $CH_3C(=O){-}NH_2$



OR



STRENGTH OF HYDROGEN BOND

- Influenced by:

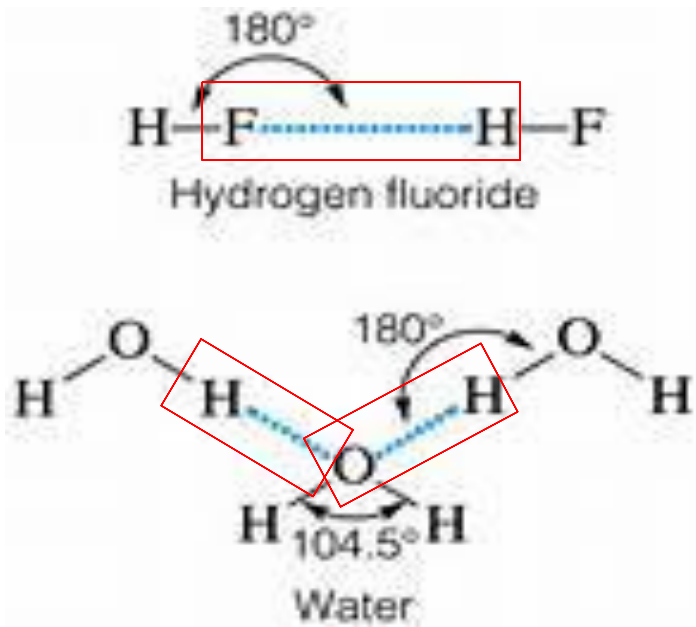
(a) Electronegativity of the element :

N < O < F

- Hydrogen bonding in H_2O is stronger than in NH_3 because O is more electronegative than N

(b) Number of hydrogen bonds

- Although F is more electronegative than O, the boiling point of HF is lower than H_2O because a molecules of H_2O can form more hydrogen bonds than HF

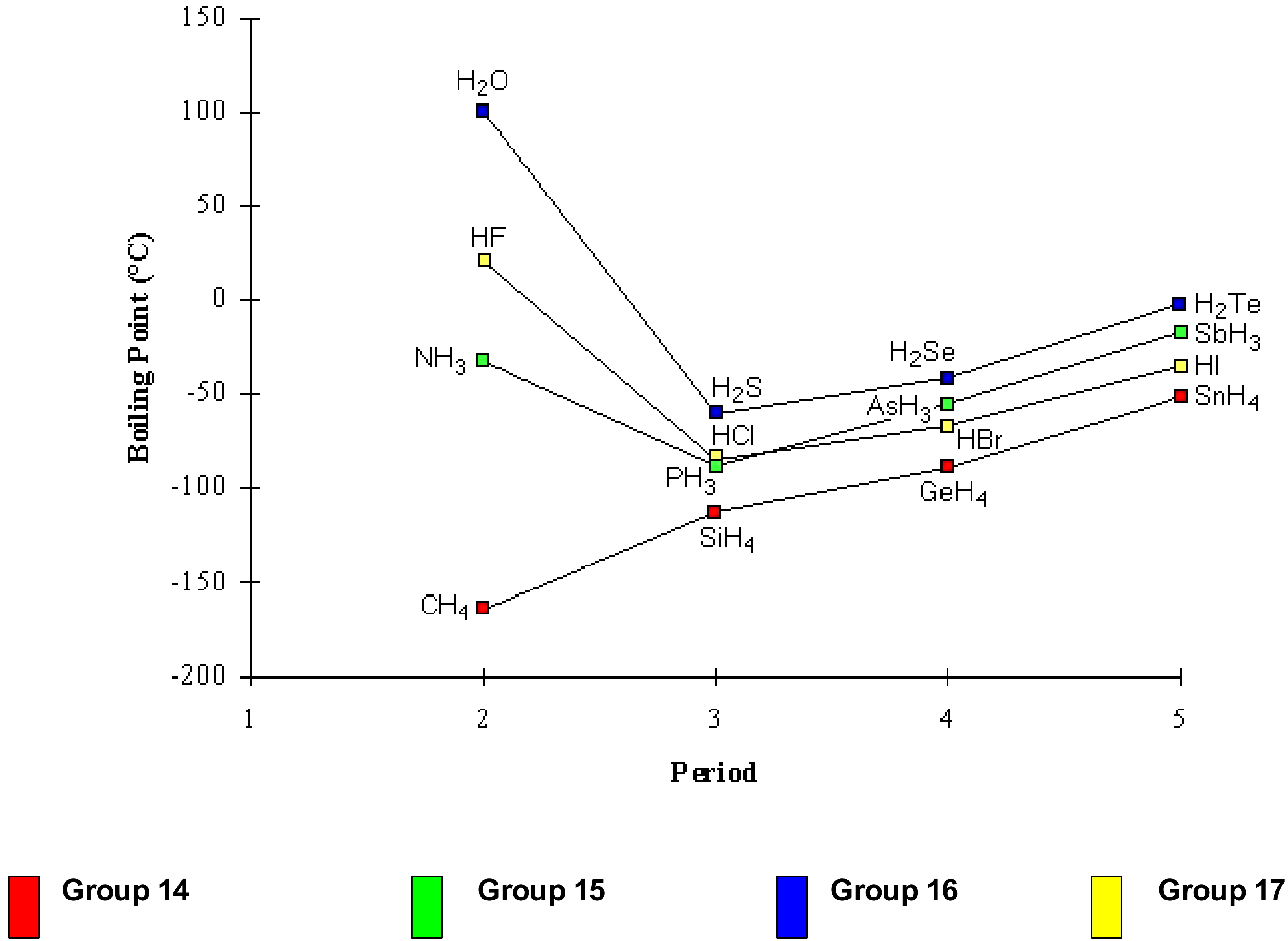


EFFECT OF HYDROGEN BOND

- **On physical properties:**
  - **Boiling point**
  - **Solubility**
  - **Density**

❖ Boiling Point

- **The boiling point of hydrides group 14, 15, 16 and 17 show the effect of hydrogen bonding on boiling points.**


150
100
50
0
-50
-100
-150
-200
Boiling Point (°C)
1
2
3
4
5
Period
H2O
HF
NH3
CH4
H2S
HCl
PH3
SiH4
H2Se
AsH3
HBr
GeH4
H2Te
SbH3
HI
SnH4
Group 14
Group 15
Group 16
Group 17

Hydrides of Group 14

- The boiling point increases gradually from CH_4, SiH_4, GeH_4 to SnH_4 because the molecular size of the hydrides increases when down the group

- The type of intermolecular forces exists in the hydrides is van der Waals forces

- The strength of van der Waals forces is proportional to the molecular weight of a substance.

Hydrides of Group 15, 16 and 17

- The boiling point for the hydrides is abnormal due to the existence of hydrogen bonding in NH_3, H_2O and HF.

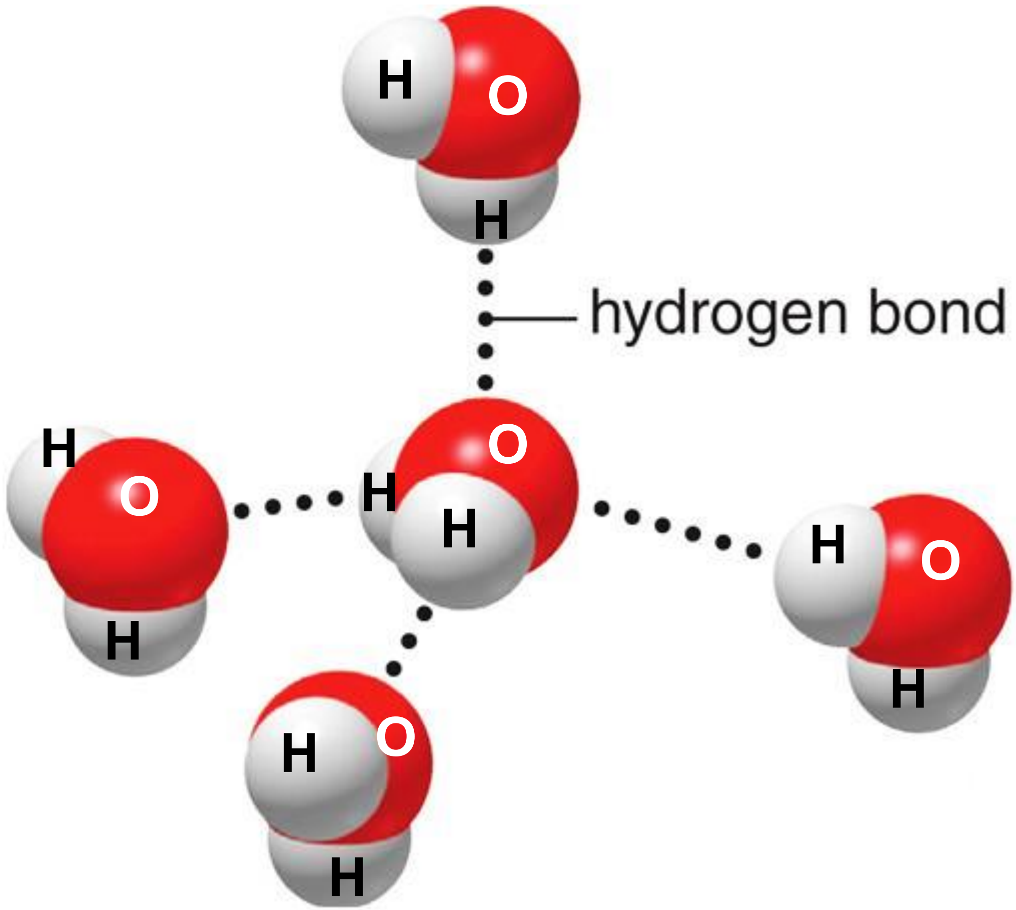
- The boiling points of NH_3, H_2O and HF are relatively higher in their respective group.

- This is due to the hydrogen bonding is stronger than van der Waals forces

- Much higher energy needed to break the hydrogen bond before the molecules can separate and enter the gas phase

Boiling point : NH_3 < HF < H_2O

- These molecules have comparable molecular weight
- F atom is more electronegative than O atom.
  The hydrogen bond between HF molecules should be stronger than H_2O.
- However, H_2O can form more hydrogen bonds per molecule. Thus, H_2O has higher boiling point than HF.
- Hydrogen bonding in NH_3 molecules is weaker than in HF molecules because N atom is less electronegative than F atom. Thus, NH_3 has lower boiling point than HF.



Water can form more hydrogen bonds per molecule.

- The increase in boiling points from PH_3 to SbH_3 (Group 15), H_2S to H_2Te (Group 16) and HCl to HI (Group 17) is due to the increasing molecular size when going down the group. Thus, the van der Waals forces increases when down the group.

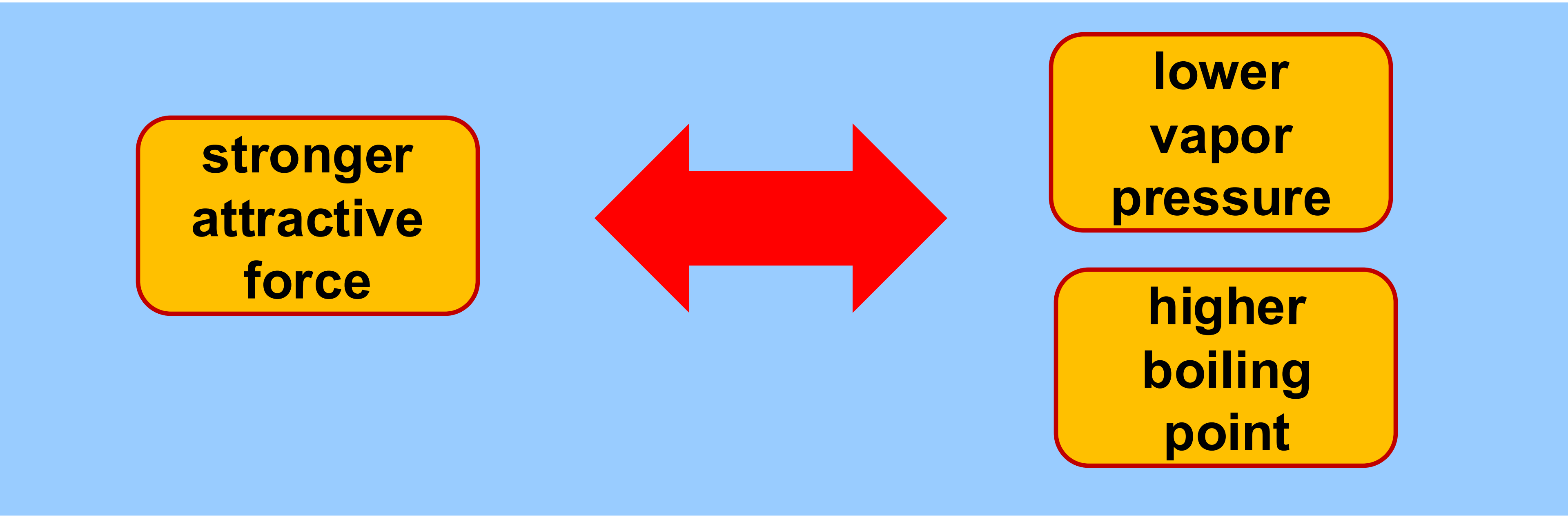
Keep in mind!

- How to compare **boiling point** between two or more substances?

Higher priority (dominant intermolecular force)

- Can **H bond** form ?
- Check the **molar mass** (London Force)
- If **molar mass** almost **same**, check the presence of **dipole–dipole force** or compare the **molecular shape**

- The stronger the **intermolecular forces**, the liquid is **less volatile** (the lower its **vapor pressure**), the higher **boiling point**

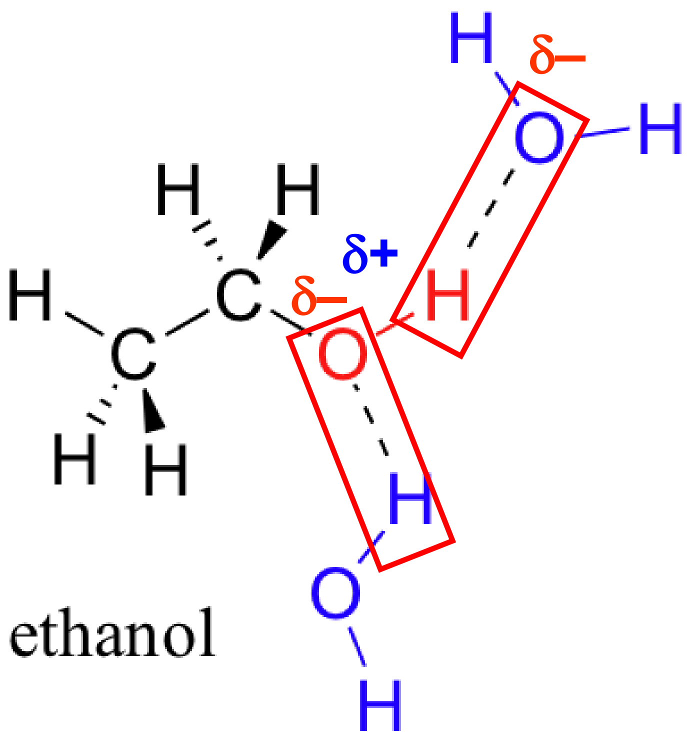


$$\text{ATTRACTIVE FORCE} \;\alpha\; \frac{\text{BOILING POINT}}{\text{VAPOR PRESSURE}}$$

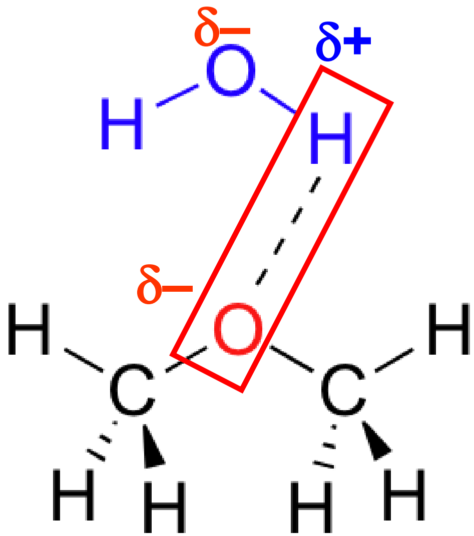
❖ Solubility

- Polar molecules tends to dissolve in polar solvents
- Water is polar and able to form hydrogen bonds
- Thus, polar molecules that can form hydrogen bond with water tends to be soluble in water

Formation of hydrogen bonding between polar molecule and water



ethanol



dimethylether

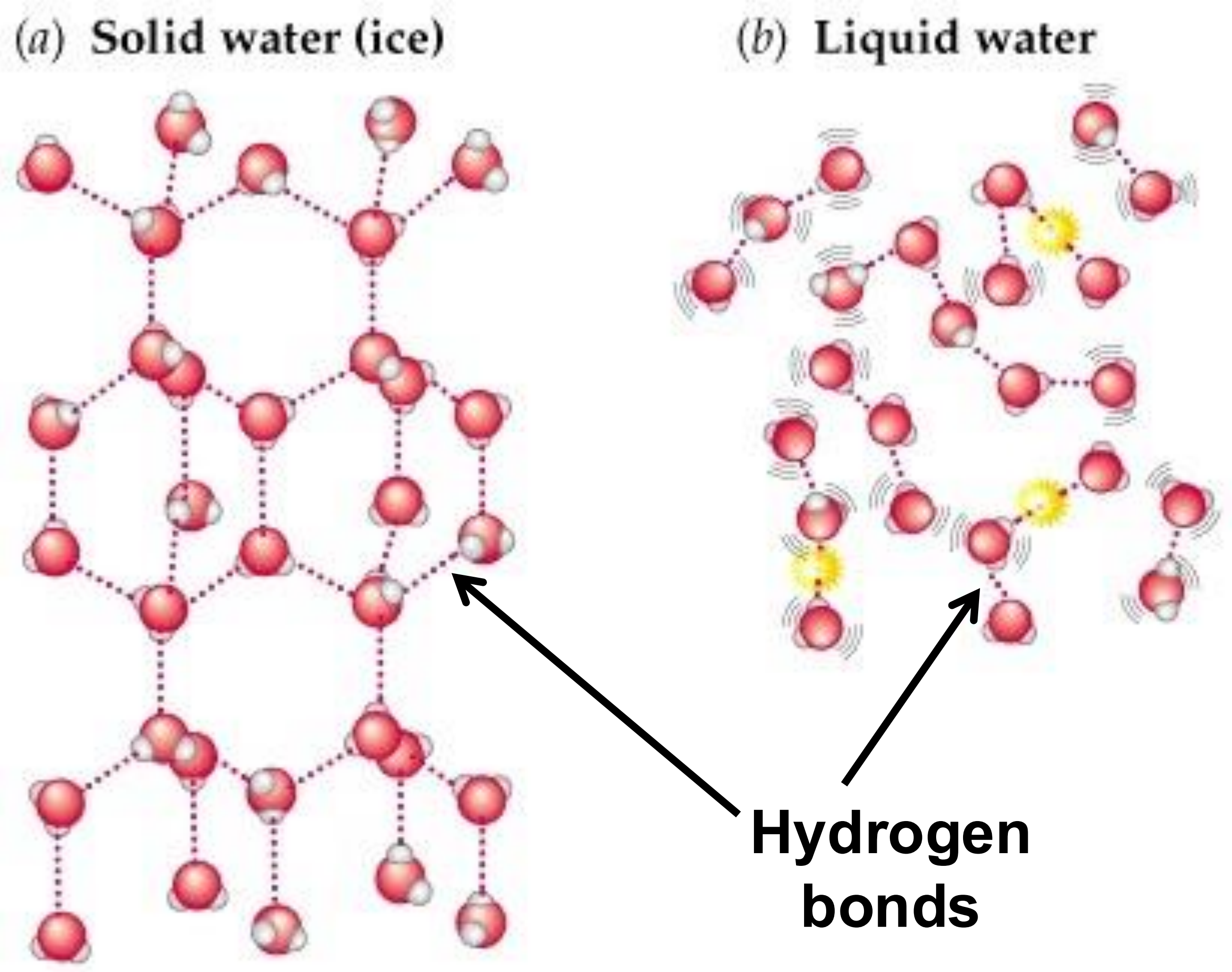
❖ Density Of Water

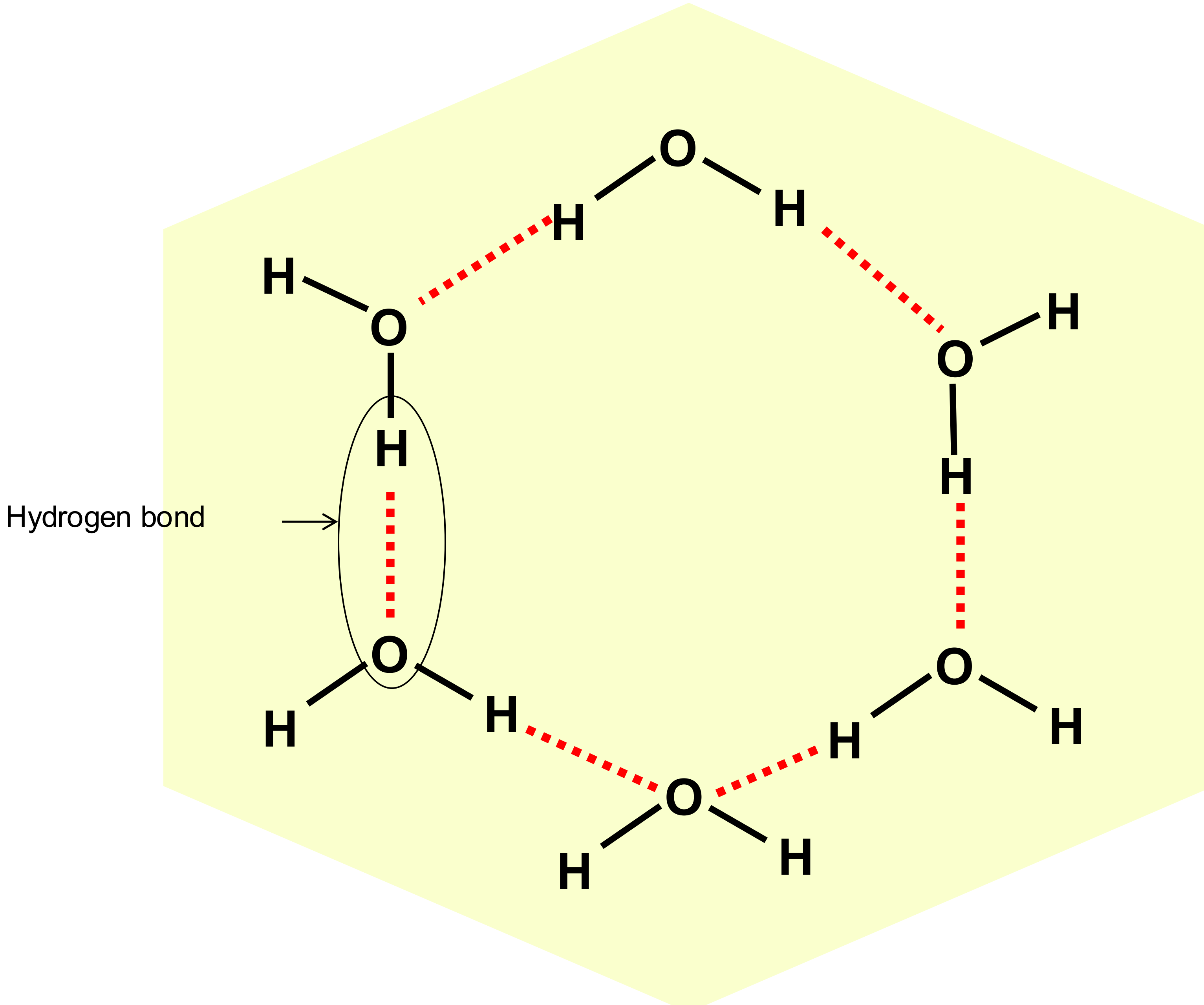
- ❑ In most substances the molecules in solid are denser than its liquid but the density of ice is less than water.

- ❑ It results from the geometrical arrangement of hydrogen bond in water. Ice has an open, hexagonally shaped crystal structure .

- ❑ When water freeze
  - ☞ volume increases
  - ☞ density decrease
  - ☞ ice floats in water

Ice crystals


(a) Solid water (ice)
(b) Liquid water
Hydrogen bonds



■ Open, hexagonally shaped ice (H_2O)

EXERCISE 1

What are strongest intermolecular force in a sample of:

(a) CH_3OH
(b) CCl_4
(c) HCl

ANSWER

a) CH_3OH

- **hydrogen bond**

(b) CCl_4

- **London forces (Dispersion forces)**

(c) HCl

- **Dipole−dipole forces**

EXERCISE – 2

Which substance has the higher boiling point?

(a) CH_3CH_2OH or $CH_3CH_2CH_3$
(b) H_2O or N_2
(c) H_2S or CH_4

ANSWER

a) CH_3CH_2OH

b) H_2O

c) H_2S

CHAPTER 4.5 : OVERVIEW

- **Formation of metallic bond**
  - Electron sea model
- **Properties of metal**
  - malleability
  - ductility
  - Electrical conductivity
  - Thermal conductivity
- **Electrical conductivity of metal**
  - Band Theory
- **Factors affecting metallic bond**
  - No. of valence electron
  - Size of atom
- **BP/MP,bonding,IMF**
  - Period 3, Group 1, Group 17

4.5 METALLIC BOND

Teaching and learning outcomes

At the end of the lesson, student should be able to

4.5	**Metallic Bond**
a)	Explain the formation of metallic bond by using electron sea model. (C2, C3)
b)	Relate metallic bond to the properties of metal: (C2, C3) i. Malleability ii. Ductility iii. Electrical conductivity iv. Thermal conductivity

4.5 METALLIC BOND

Teaching and learning outcomes

At the end of the lesson, student should be able to

4.5 Metallic Bond

c) Explain electrical conductivity of metal by using band theory. (C2, C3)

d) Explain the factors that affect the strength of metallic bond. (C2, C3)

e) Relate boiling/melting point to the molecular structure, types of bonding and intermolecular forces for element of (C2, C3)

   i. Period 3

   ii. Group 1

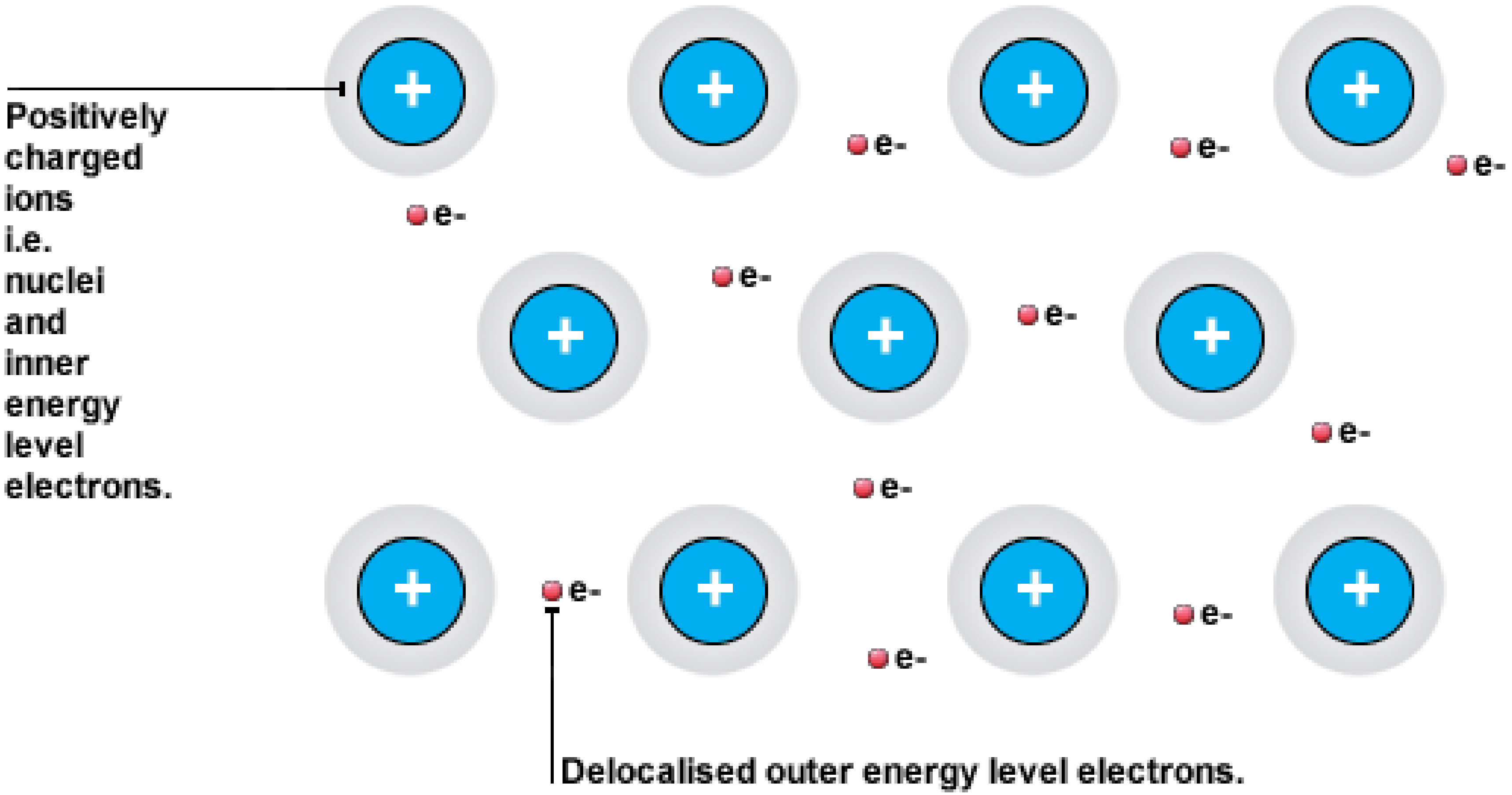
   iii. Group 17

METALLIC BOND

- **Electrostatic attraction between the positively charged metal ions and the "sea of delocalised "valence electrons**

- **The electron sea model is used to explain the metallic bond exist in metallic elements such as Na, Mg and Al**

"ELECTRON–SEA" MODEL



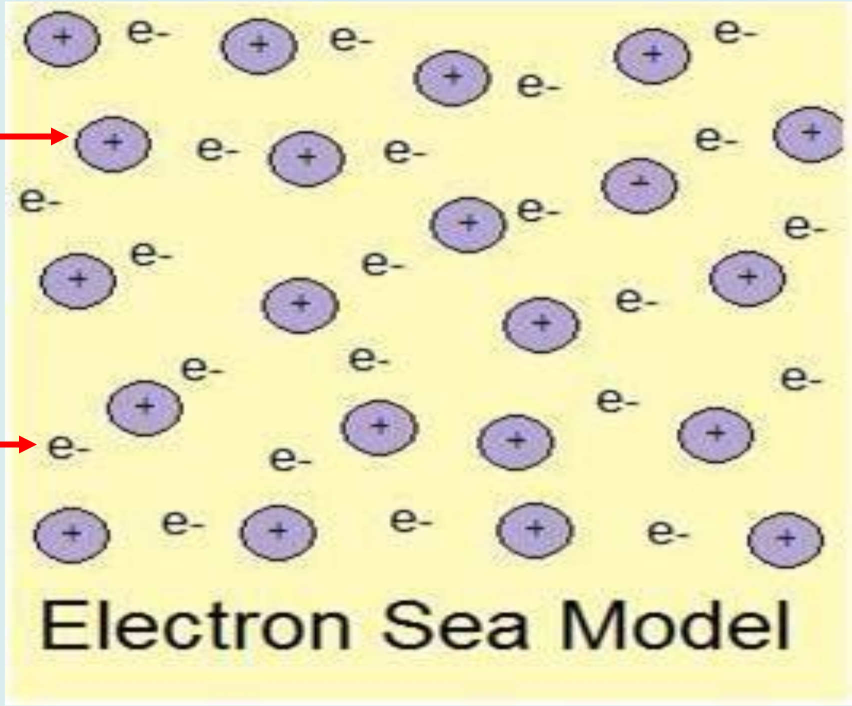
- **In the solid state, the metal atoms are packed as close as possible.**

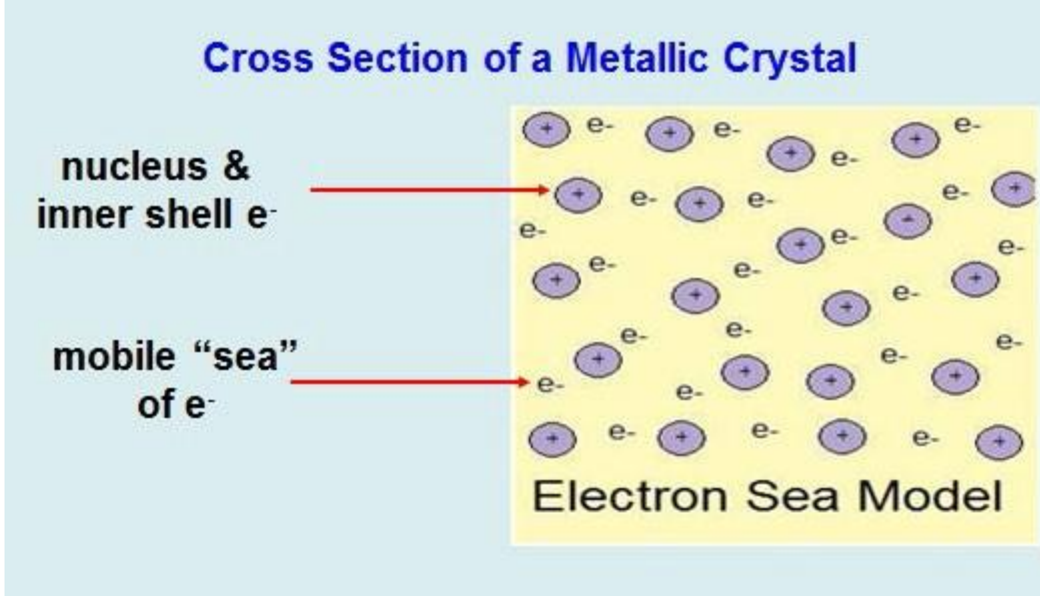
- **All metal atoms in the sample contribute their valence electrons to form an "electron sea" that is delocalized over the entire solid.**

Cross Section of a Metallic Crystal

nucleus & inner shell e^-

mobile "sea" of e^-



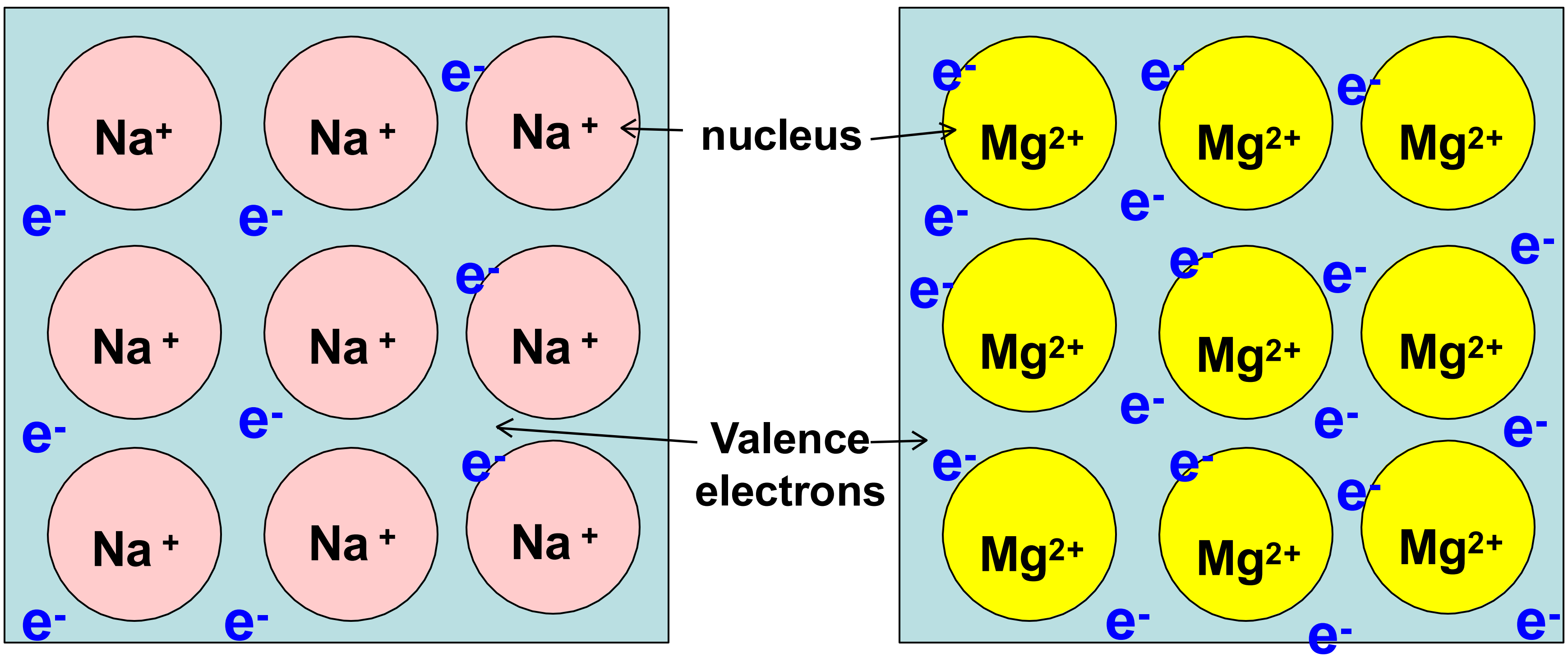


- Metallic bond formed from the electrostatic attraction between the positively charged metal ions (nuclei) and the "sea of delocalized" valence electrons.

- The sea of valence electrons is acting as 'glue' bonding the positive ions (which would otherwise repel each other)

Example: “Electron-sea” Model

Group 1 | Group 2



Sodium, Na | Magnesium, Mg

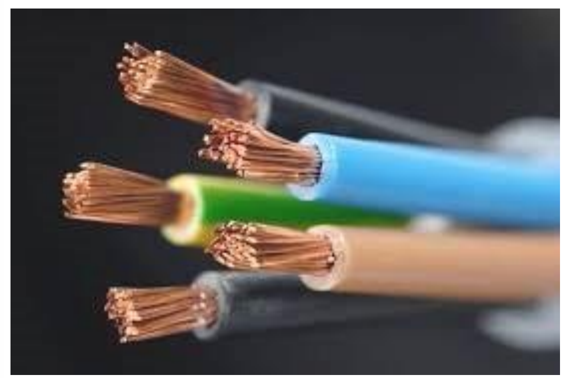
PROPERTIES OF METALS

- **Lustrous** in appearance

- Most are solids with moderate to high melting point and much higher boiling point.
- Bent or dent rather than crack or shatter

- **Malleable** □ **flattened into sheet**


- **Ductile** □ **pulled into wires**

- **Conduct heat and electricity well in both the solid and liquid states**

Aluminum foil

A sheet of aluminum foil is made up of metallic bonds.

Why most are solids with moderate melting point and much higher boiling point?

❑ Explanation:

- Melting point is moderate because the attraction between moveable cations (nuclei) and mobile e^- doesn't need to be broken during melting

- Boiling point is higher because it requires each cation (nuclei) and the mobile e^- to break away from the others

□ the metallic bond is strong enough to resist separation of the atom

Why metals are malleable and ductile?

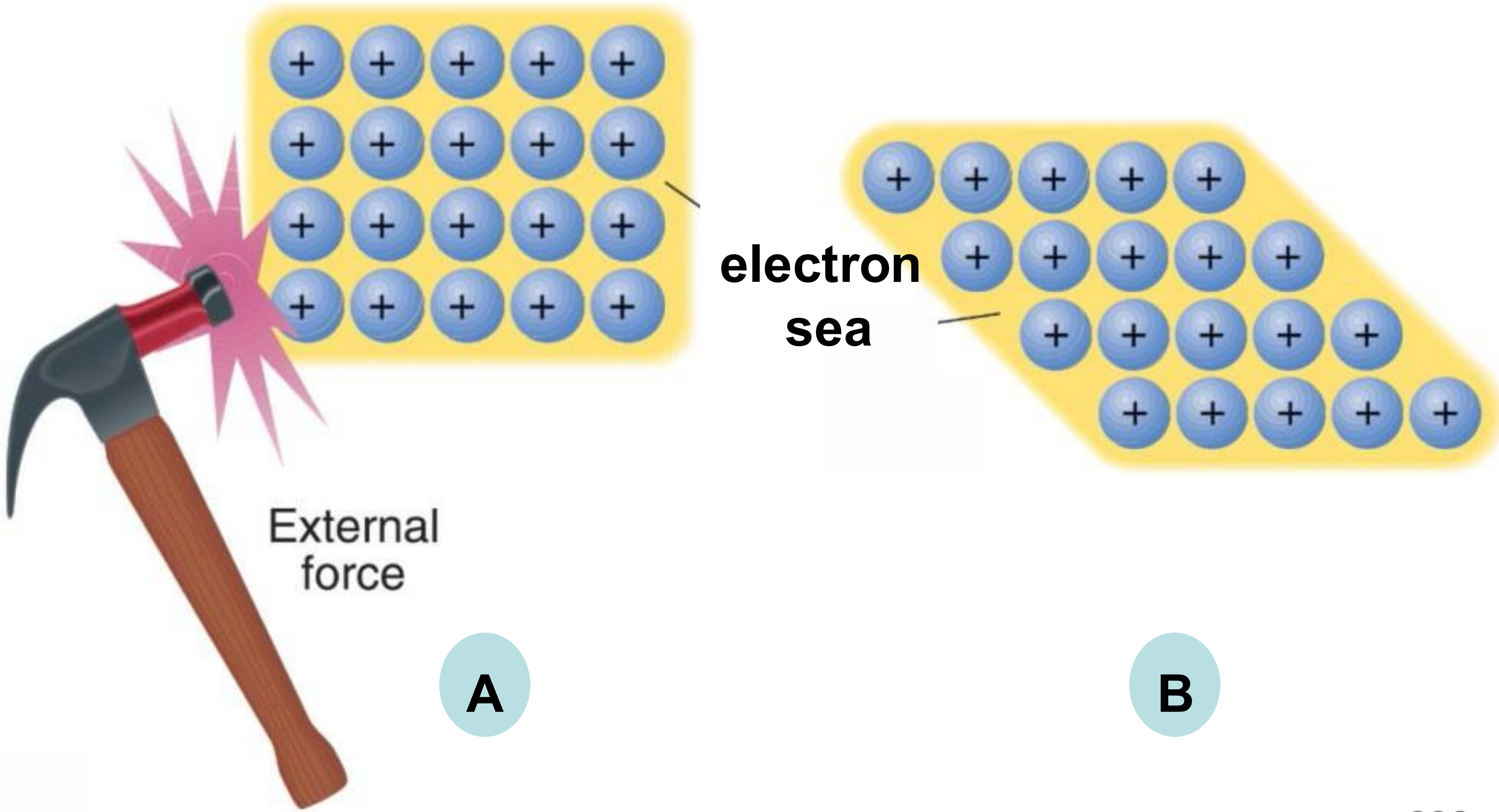
- **Explanation:**

  - **Metal atoms form metallic bonds to many neighbors.**

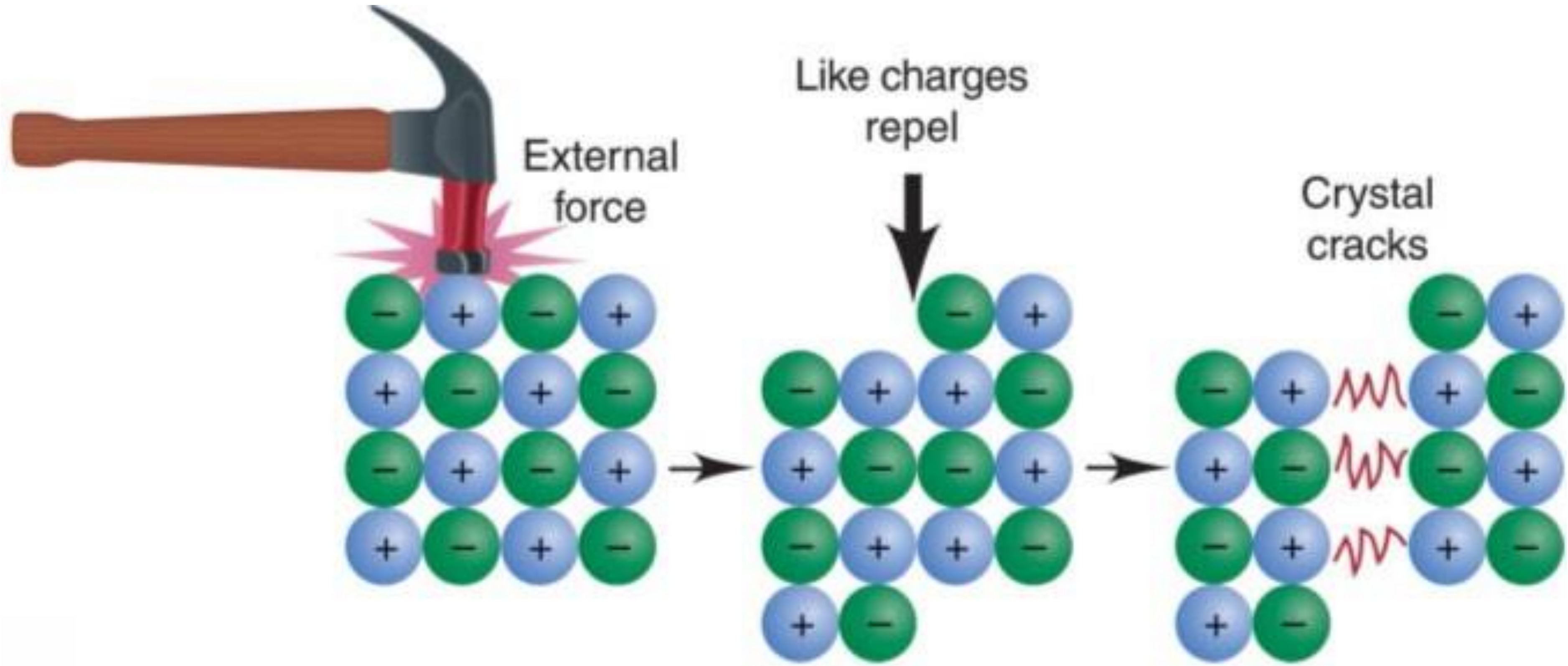
  - **When a piece of metal deformed by a hammer, the metal ions (nuclei) slide past each other through the e^- sea to new positions.**

- **Changes in the positions of the metal ions brought about reshaping the metal accommodated by redistribution of valence elecrons.**

The Metal Ions (Nuclei) slide past each



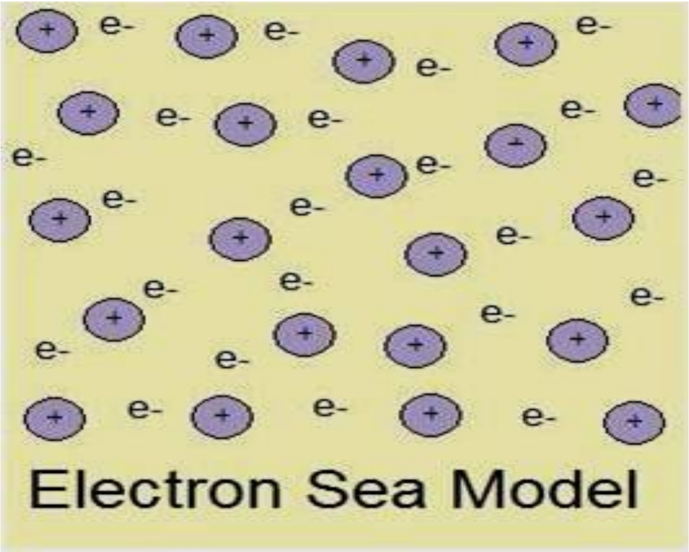
Ionic Compounds crack due to Electrostatic Force



Why metal conduct heat and electricity well in both the solid and liquid states ?

- Explanation:

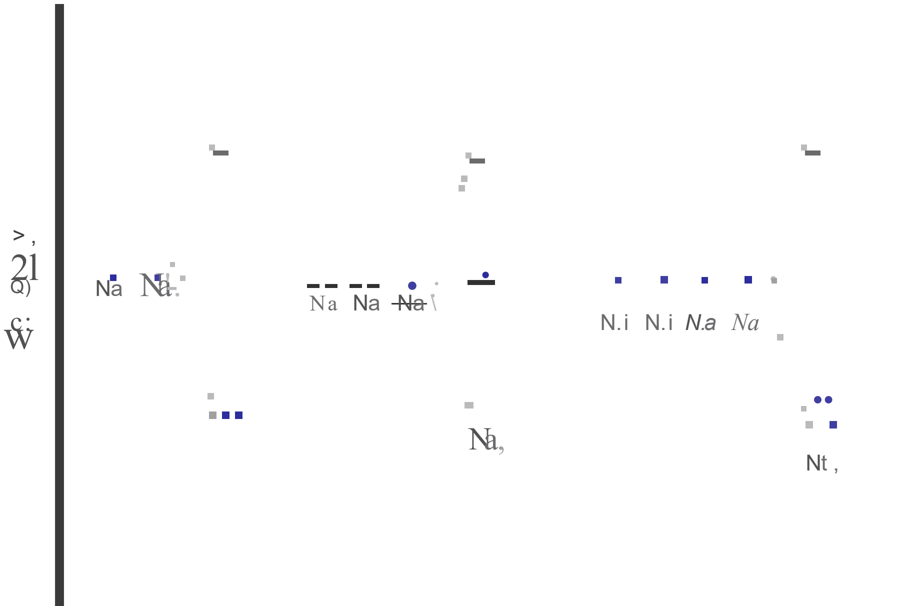
- Because metals atoms have mobile e^-

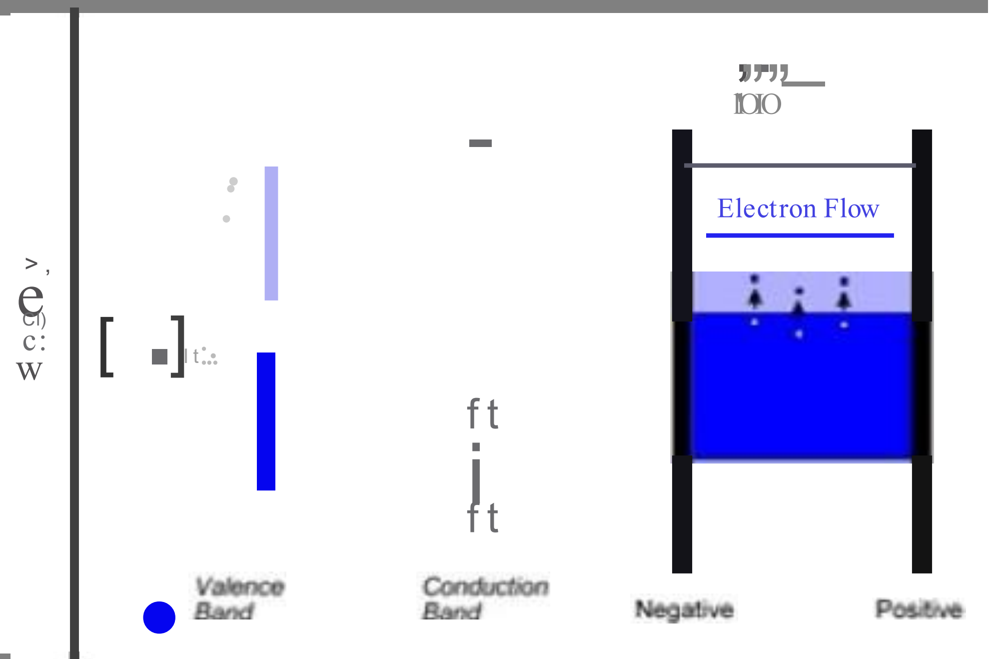


BAND THEORY OF ELECTRICAL CONDUCTIVITY

Band theory of solids

- The delocalized electrons move freely through "band" formed by overlapping molecular orbitals.
- The electronic structure of a bulk solid is referred to as a band structure
- When valence band (lower energy) and conduction band (higher energy) overlapping, allowing electrons to flow through the metals with minimal applied voltage.




Electron Flow
Valence Band
Conduction Band
Negative
Positive

BAND THEORY

- **Band**: An array of closely spaced molecular orbitals occupying a continuous range of energy
- **Band gap**: The energy gap between a fully occupied valence band and an empty conduction band.
- **Conduction band**: A band of unoccupied molecular orbitals lying higher in energy than the occupied valence band
- **Valence band**: a band closely spaced bonding molecular orbitals that is essentially fully occupied by electron

BAND THEORY

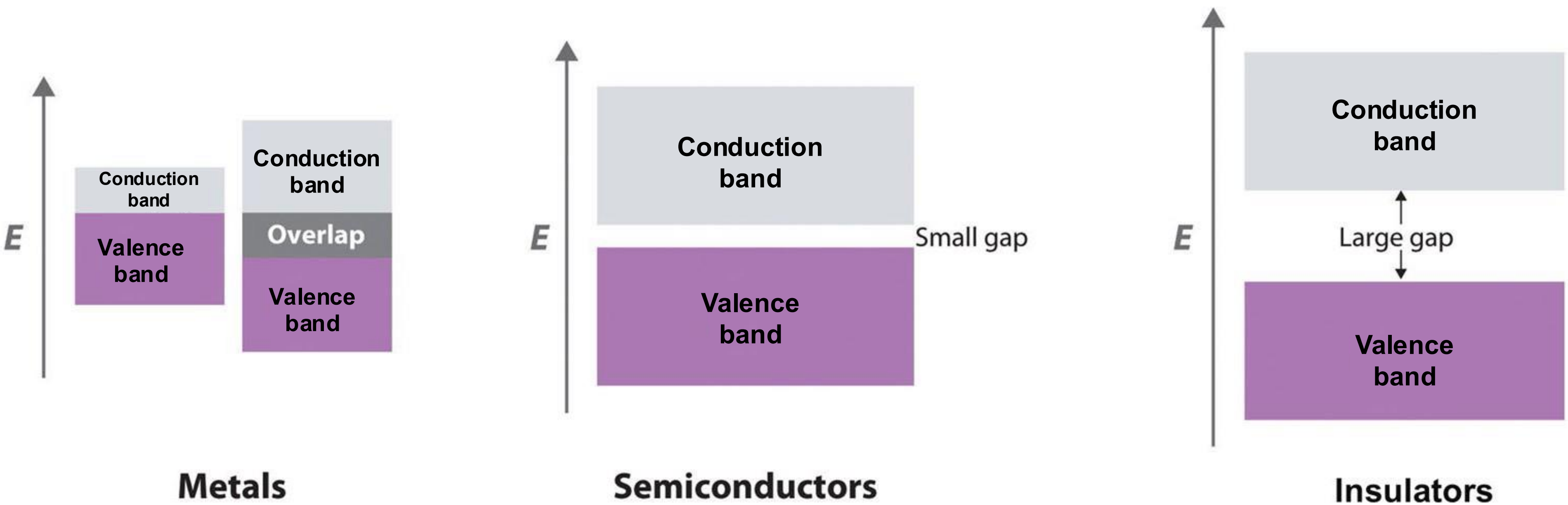
Band theory of solids:

- **electrons jump from valence band to conduction band even at ordinary temperature and if this happens then the solid conducts electricity.**

- **Conductivity depends on the gap between the valence band and conduction band.**

Conductors	Semiconductors	Insulators
There is no band gap between their valence band and conduction bands, since they overlap. There is a continuous availability of electrons in these closely spaced orbitals.	have a small energy gap between the valence band and the conduction band. Electrons can make the jump up to the conduction band, but not with the same ease as they do in conductors.	The band gap between the valence band the conduction band is so large that electrons cannot make the energy jump from the valence band to the conduction band.

Band structure of conductors, semiconductors and insulators

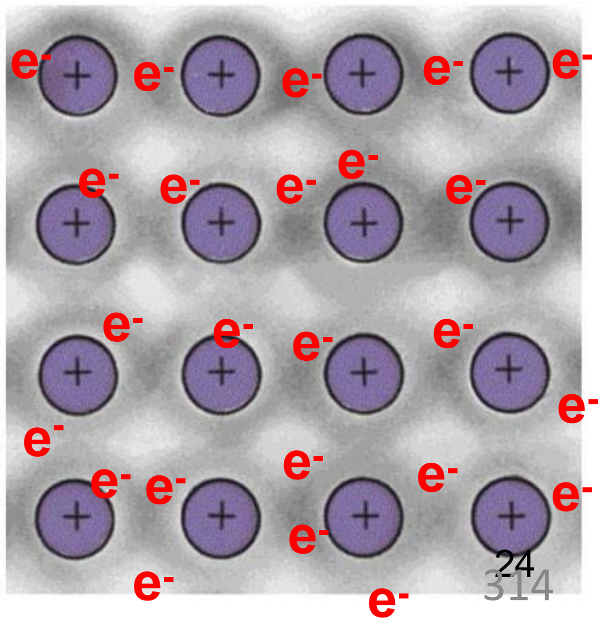


STRENGTH OF METALLIC BOND

- Factor affecting:

  1. Number of valence e^-

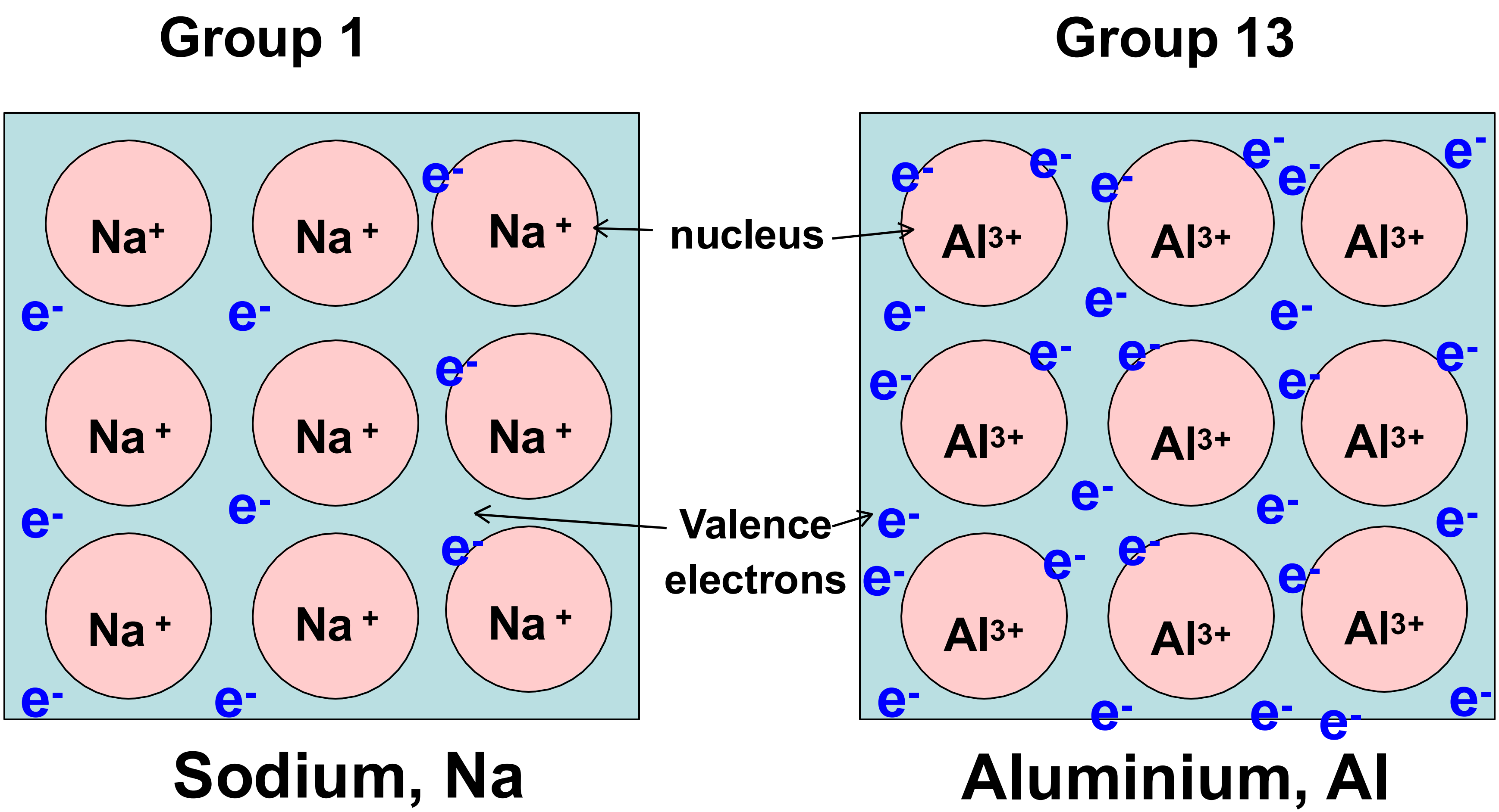
  2. Size of atoms:



❶ Number of valence e^- :

- More valence e^- , more delocalized e^-
- Attraction between positive ions (nuclei) and delocalized e^- stronger
- Stronger metallic bond

Example:



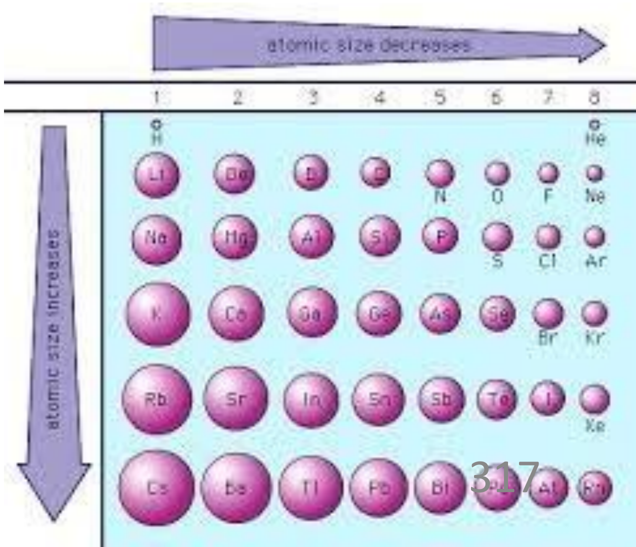
- ❑ Metallic bond of Aluminium, Al stronger than Sodium, Na

❷ Size of atoms :

- **Smaller size**

- **Attraction between positive ions (nuclei) and delocalized e^- stronger**

- **Stronger metallic bond**

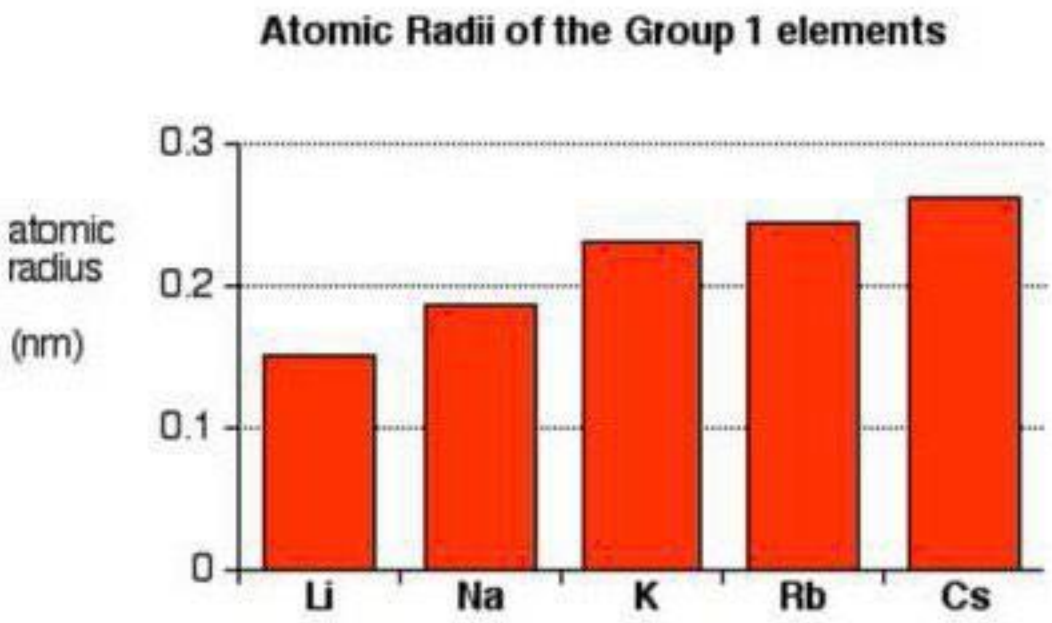


Across Period 3:

Element	Na	Mg	Al
No. of valence e^- per atom	1	2	3
M.P (^oC)	97.8	651	660
B.P (^oC)	892	1107	2467

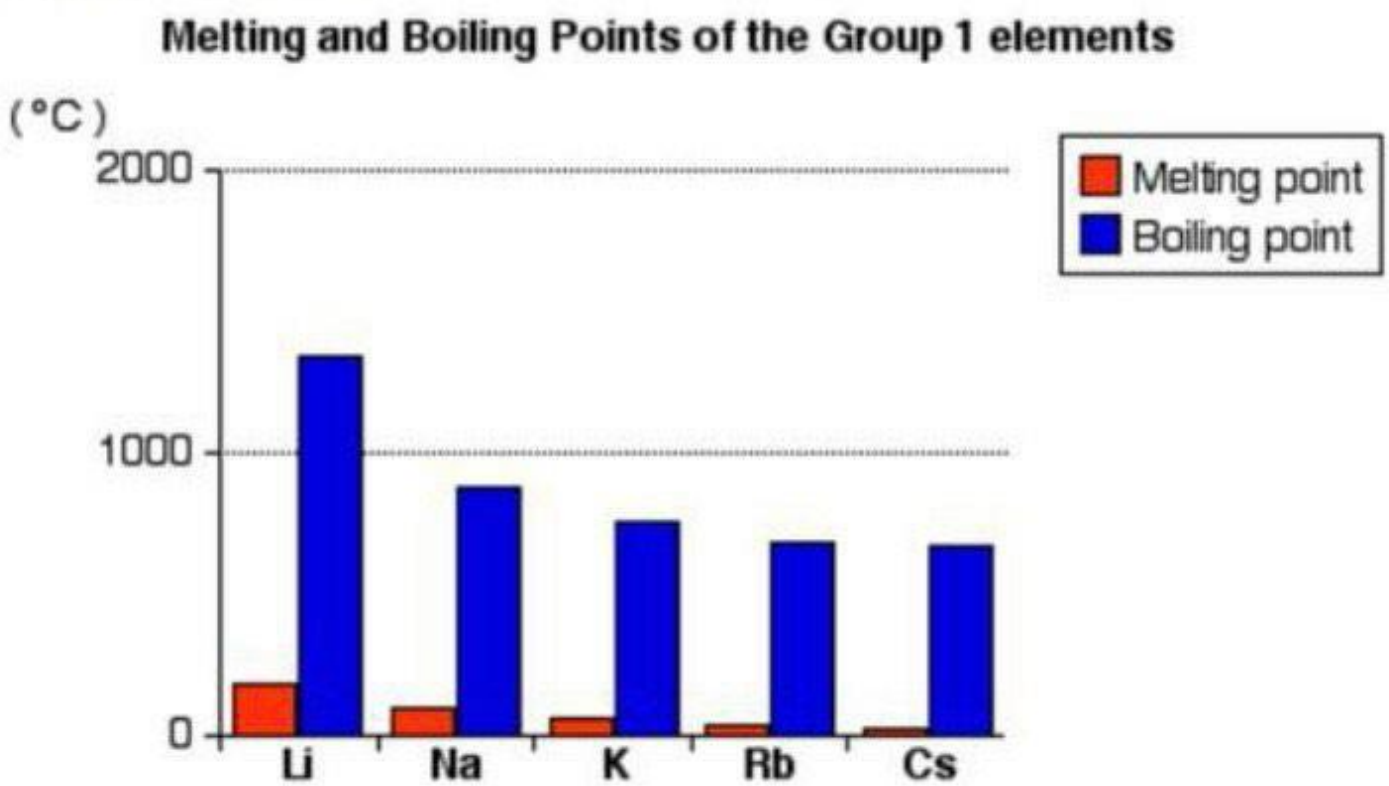
- More valence e^- , more delocalized e^-
- Attraction between positive ions (nuclei) and delocalized e^- stronger
- Stronger metallic bond, higher boiling point

- D g



- **The atoms in a metal are held together by the attraction of the nuclei to electrons which are delocalized over the whole metal.**

- **As the atoms increase in size, the distance between nuclei and these delocalized electrons increases, therefore metallic bond of the atom is getting weaker.**



- The **decrease in melting and boiling points reflects the decrease in the strength of each metallic bond.**

Down group 17:

Halogen	Melting point (°C)	Boiling point (°C)
F_2	-220.0	-188.0
Cl_2	-101.0	-35.0
Br_2	-7.2	58.8
I_2	114.0	184.0

- The melting and boiling point increase down the group because of the van der Waals forces.
- The size of the molecules increases down the group.
- This increase in size means an increase in the strength of the van der Waals forces.

Thanks!

For Attention

See You The Next Chapter

End Slide

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