

# Formule di Lewis, Risonanza, Carica Formale e VSEPR

## Configurazioni elettroniche esterne e elettroni di valenza

**Configurazione elettronica esterna: solo gli elettroni nel livello energetico più alto (valore di  $n$  più alto).**

Gli elementi nello stesso gruppo del sistema periodico hanno la stessa configurazione elettronica esterna.

***Configurazioni elettroniche simili sono correlate con comportamenti chimici simili.***



**elettroni di valenza** sono quelli che intervengono nella formazione dei composti.

# Configurazioni elettroniche condensate dello stato fondamentale nei primi tre periodi

|   | 1A<br>(1)                      | 2A<br>(2)                      | 3A<br>(13)                          | 4A<br>(14)                          | 5A<br>(15)                         | 6A<br>(16)                         | 7A<br>(17)                          | 8A<br>(18)                          |
|---|--------------------------------|--------------------------------|-------------------------------------|-------------------------------------|------------------------------------|------------------------------------|-------------------------------------|-------------------------------------|
| 1 | 1<br>H<br>$1s^1$               |                                |                                     |                                     |                                    |                                    |                                     | 2<br>He<br>$1s^2$                   |
| 2 | 3<br>Li<br>$[\text{He}] 2s^1$  | 4<br>Be<br>$[\text{He}] 2s^2$  | 5<br>B<br>$[\text{He}] 2s^2 2p^1$   | 6<br>C<br>$[\text{He}] 2s^2 2p^2$   | 7<br>N<br>$[\text{He}] 2s^2 2p^3$  | 8<br>O<br>$[\text{He}] 2s^2 2p^4$  | 9<br>F<br>$[\text{He}] 2s^2 2p^5$   | 10<br>Ne<br>$[\text{He}] 2s^2 2p^6$ |
| 3 | 11<br>Na<br>$[\text{Ne}] 3s^1$ | 12<br>Mg<br>$[\text{Ne}] 3s^2$ | 13<br>Al<br>$[\text{Ne}] 3s^2 3p^1$ | 14<br>Si<br>$[\text{Ne}] 3s^2 3p^2$ | 15<br>P<br>$[\text{Ne}] 3s^2 3p^3$ | 16<br>S<br>$[\text{Ne}] 3s^2 3p^4$ | 17<br>Cl<br>$[\text{Ne}] 3s^2 3p^5$ | 18<br>Ar<br>$[\text{Ne}] 3s^2 3p^6$ |

# Simboli di Lewis

Per scrivere il simbolo di Lewis di un elemento dei gruppi principali:

- Notare il **numero del Gruppo A** che dà il numero di elettroni di valenza.
- Collocare un **puntino sui quattro lati del simbolo** dell'elemento.
- Aggiungere **puntini appaiandoli fino a che non sono stati tutti utilizzati (Pauli)**

Esempio:

L'azoto, N, è nel Gruppo 5A e ha perciò 5 elettroni di valenza:



Desper Collection in the History of Chemistry/University of Cincinnati

**Gilbert Newton Lewis (1875–1946)** Lewis introduced the theory of shared electron-pair chemical bonds in a paper published in the *Journal of the American Chemical Society* in 1916. Lewis also made major contributions in acid–base chemistry, thermodynamics, and the interaction of light with substances.

Lewis was born in Massachusetts but raised in Nebraska. After earning his B.A. and Ph.D. at Harvard, he began his career in 1912 at the University of California at Berkeley. He was not only a productive researcher but was also an influential teacher. Among his ideas was the use of problem sets in teaching, a practice still in use today.

# Simboli di Lewis per gli elementi dei Periodi 2 e 3

Periodo

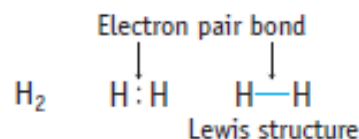
|   | 1A(1)  | 2A(2)  |
|---|--------|--------|
|   | $ns^1$ | $ns^2$ |
| 2 | • Li   | • Be • |
| 3 | • Na   | • Mg • |

| 3A(13)     | 4A(14)     | 5A(15)     | 6A(16)     | 7A(17)     | 8A(18)     |
|------------|------------|------------|------------|------------|------------|
| $ns^2np^1$ | $ns^2np^2$ | $ns^2np^3$ | $ns^2np^4$ | $ns^2np^5$ | $ns^2np^6$ |
| • B •      | • C •      | • N •      | • O •      | • F •      | • Ne •     |
| • Al •     | • Si •     | • P •      | • S •      | • Cl •     | • Ar •     |

# Simboli di Lewis e legami

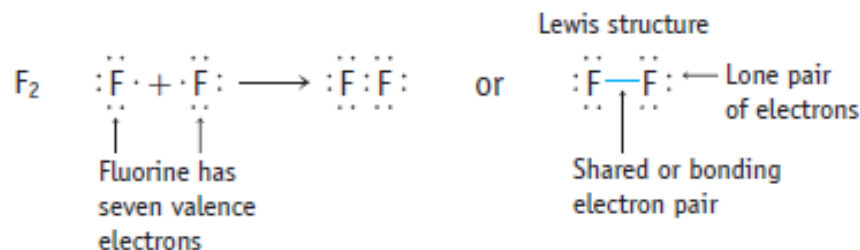
The electron pair bond between the two atoms of an  $H_2$  molecule is represented by a pair of dots or, alternatively, a line.

The representation of a molecule in this fashion is called a **Lewis electron dot structure** or just a **Lewis structure**.



Simple Lewis structures, such as that for  $F_2$ , can be drawn starting with Lewis dot symbols for atoms and arranging the valence electrons to form bonds. Fluorine, an element in Group 7A, has seven valence electrons. The Lewis symbol shows that an F atom has a single unpaired electron along with three electron pairs. In  $F_2$ , the single electrons, one on each F atom, pair up in the covalent bond.

In the Lewis structure for  $F_2$  the pair of electrons in the F—F bond is the bonding pair, or **bond pair**. The other six pairs reside on single atoms and are called **lone pairs**. Because they are not involved in bonding, they are also called **nonbonding electrons**.



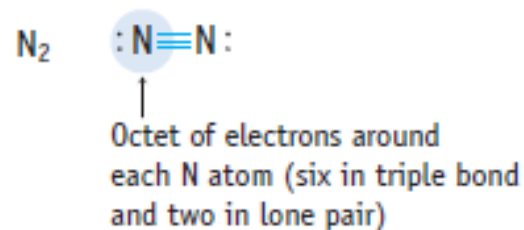
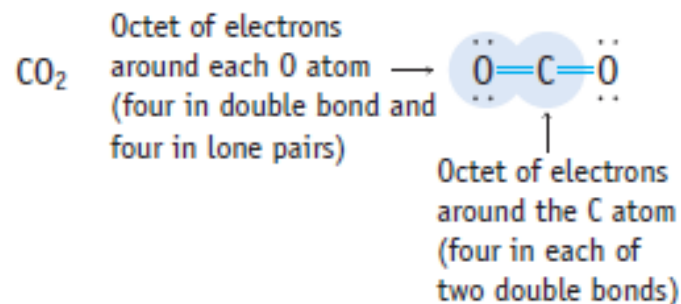
**Regola dell'ottetto:** quando gli atomi si legano, essi cedono, acquistano o condividono elettroni per raggiungere un livello esterno pieno di 8 (2 per H e Li) elettroni.

# Formule di Lewis per molecole

Carbon dioxide,  $\text{CO}_2$ , and dinitrogen,  $\text{N}_2$ , are examples of molecules in which two atoms share more than one electron pair.

In carbon dioxide, the carbon atom shares two pairs of electrons with each oxygen and so is linked to each O atom by a **double bond**. The valence shell of each oxygen atom in  $\text{CO}_2$  has two bonding pairs and two lone pairs.

In dinitrogen, the two nitrogen atoms share three pairs of electrons, so they are linked by a **triple bond**. In addition, each N atom has a single lone pair.

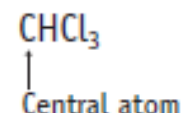


# Formule di Lewis per molecole

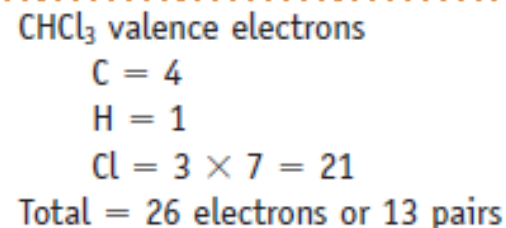
## Drawing Lewis Electron Dot Structures

A systematic approach to constructing Lewis structures is illustrated with chloroform,  $\text{CHCl}_3$ , a compound once used as an anesthetic.

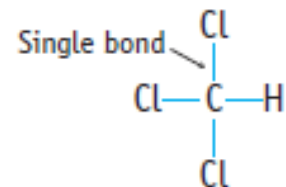
**STEP 1** Determine the arrangement of atoms within a molecule. The central atom is *usually* the one with the lowest affinity for electrons (least negative electron attachment enthalpy, Section 7-5).



**STEP 2** Determine the total number of valence electrons in the molecule or ion. In a neutral molecule, this number will be the sum of the valence electrons for each atom. The number of valence electron *pairs* will be half the total number of valence electrons.



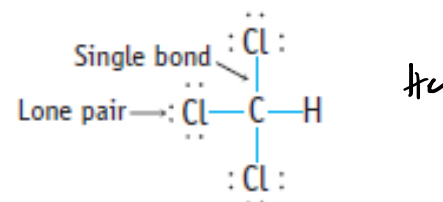
**STEP 3** Place one pair of electrons between each pair of bonded atoms to form a single bond. Here, four electron pairs are used to make four single bonds (which are represented by single lines). Nine pairs of electrons, of the original 13, remain to be used.





**STEP 4** Use remaining pairs as lone pairs around each terminal atom (except H) so that each terminal atom is surrounded by eight electrons. If, after this is done, there are electrons left over, assign them to the central atom. All 13 pairs have now been assigned in  $\text{CHCl}_3$ . The H atom has a share in two electrons, as it normally does, and all other atoms have a share in an octet of electrons. The Lewis structure for  $\text{CHCl}_3$  is complete.

$$26 - 8 = 18$$



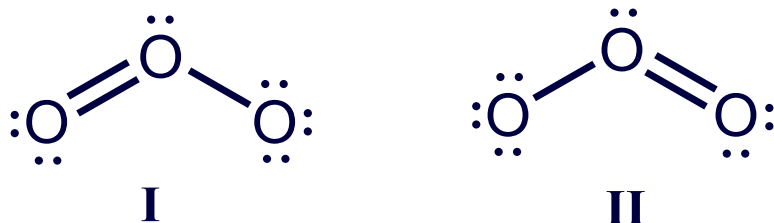
**STEP 5** If no valence electron pairs remain after forming single bonds and completing the octets of terminal atoms, and the central atom does not have an octet of electrons, then multiple bonds can be created by sharing one or more pairs of electrons between terminal atoms and the central atom. (See the case of  $\text{CH}_2\text{O}$  on page 289).

This step is not required for  $\text{CHCl}_3$ .

| Group 4A                            |  | Group 5A                            |  | Group 6A                                    |  | Group 7A                         |                     |
|-------------------------------------|--|-------------------------------------|--|---|--|----------------------------------|---------------------|
| $\text{CH}_4$<br>methane            | $\begin{array}{c} \text{H} \\   \\ \text{H}-\text{C}-\text{H} \\   \\ \text{H} \end{array}$  | $\text{NH}_3$<br>ammonia            | $\begin{array}{c} \text{H} \\   \\ \text{H}-\text{N}-\text{H} \\   \\ \text{H} \end{array}$  | $\text{H}_2\text{O}$<br>water               | $\begin{array}{c} \text{H} \\   \\ \text{H}-\text{O}-\text{H} \\   \\ \text{H} \end{array}$  | $\text{HF}$<br>hydrogen fluoride | $\text{H}-\text{F}$ |
| $\text{C}_2\text{H}_6$<br>ethane    | $\begin{array}{c} \text{H} \quad \text{H} \\   \quad   \\ \text{H}-\text{C}-\text{C}-\text{H} \\   \quad   \\ \text{H} \quad \text{H} \end{array}$ | $\text{N}_2\text{H}_4$<br>hydrazine | $\begin{array}{c} \text{H} \quad \text{H} \\   \quad   \\ \text{H}-\text{N}-\text{N}-\text{H} \\   \quad   \\ \text{H} \quad \text{H} \end{array}$ | $\text{H}_2\text{O}_2$<br>hydrogen peroxide | $\begin{array}{c} \text{H} \quad \text{H} \\   \quad   \\ \text{H}-\text{O}-\text{O}-\text{H} \\   \quad   \\ \text{H} \quad \text{H} \end{array}$ |                                  |                     |
| $\text{C}_2\text{H}_4$<br>ethylene  | $\begin{array}{c} \text{H} \quad \text{H} \\   \quad   \\ \text{H}-\text{C}=\text{C}-\text{H} \\   \quad   \\ \text{H} \quad \text{H} \end{array}$ | $\text{NH}_4^+$<br>ammonium ion     | $\left[ \begin{array}{c} \text{H} \\   \\ \text{H}-\text{N}-\text{H} \\   \\ \text{H} \end{array} \right]^+$                                       | $\text{H}_3\text{O}^+$<br>hydronium ion     | $\left[ \begin{array}{c} \text{H} \\   \\ \text{H}-\text{O}-\text{H} \\   \\ \text{H} \end{array} \right]^+$                                       |                                  |                     |
| $\text{C}_2\text{H}_2$<br>acetylene | $\text{H}-\text{C}\equiv\text{C}-\text{H}$   | $\text{NH}_2^-$<br>amide ion        | $\left[ \begin{array}{c} \text{H} \\   \\ \text{H}-\text{N}-\text{H} \\   \\ \text{H} \end{array} \right]^-$                                       | $\text{OH}^-$<br>hydroxide ion              | $\left[ \begin{array}{c} \text{H} \\   \\ \text{O}-\text{H} \\   \\ \text{H} \end{array} \right]^-$  |                                  |                     |

## Strutture di risonanza

O<sub>3</sub> può essere descritto da 2 strutture di Lewis:

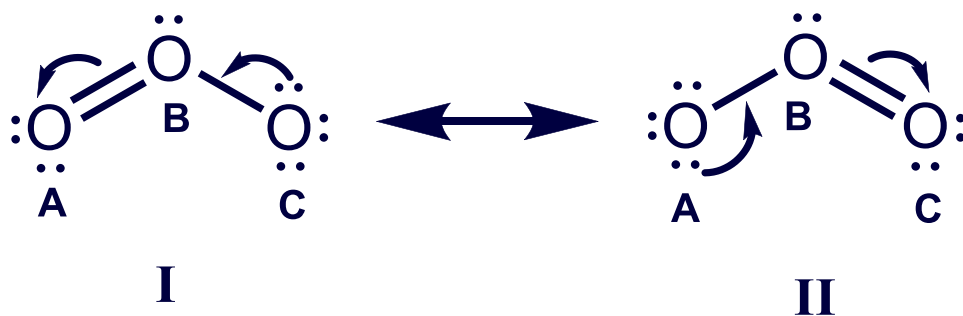


Entrambe le strutture sono corrette e descrivono la **stessa** molecola.

**Nessuna** delle due strutture di Lewis rappresenta accuratamente O<sub>3</sub>. Le misure delle lunghezze di legame e delle energie di legame indicano che i due legami in O<sub>3</sub> sono identici, con proprietà che si situano tra quelle di un legame semplice e quelle di un legame doppio.

## Strutture di risonanza

La struttura di  $O_3$  è rappresentata più correttamente con due strutture di Lewis, dette **strutture di risonanza** (o **forme di risonanza**)



Le strutture di risonanza *hanno la stessa posizione relativa degli atomi ma differenti posizioni delle coppie di elettroni di legame e di elettroni solitari.*

## L'ibrido di risonanza

Una specie come  $O_3$ , che può essere rappresentata da più di una formula di Lewis valida, è detta un ***ibrido di risonanza***.

Le forme di risonanza ***non rappresentano dei veri legami***.

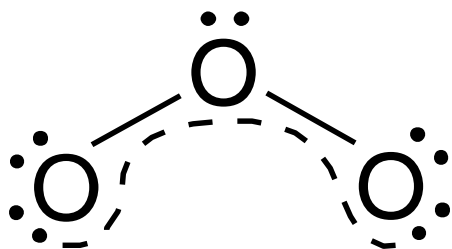
$O_3$  ***non*** si trasforma continuamente tra le sue due forme di risonanza.

La ***vera struttura*** di un ibrido di risonanza è una ***media*** delle sue forme di risonanza

# Delocalizzazione degli elettroni

Nelle strutture di Lewis gli elettroni sono **localizzati** su un singolo atomo (lone pairs) o in un legame tra due atomi (coppia condivisa).

In un ibrido di risonanza, gli elettroni sono **delocalizzati**: la densità elettronica è “distribuita” su più atomi adiacenti.



L'ibrido di risonanza è disegnato con una linea curva tratteggiata per rappresentare le coppie delocalizzate

**Presenza di elettroni  $\pi$  delocalizzati**

**Presenza di legami doppi adiacenti a legami singoli**



**RISONANZA**

# Carica formale e numero di ossidazione

Nella determinazione della **carica formale**, gli elettroni di legame sono **condivisi ugualmente** dagli atomi.

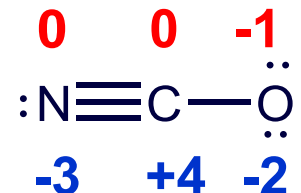
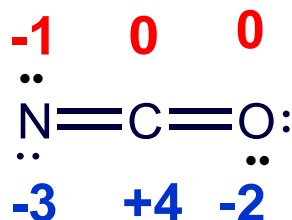
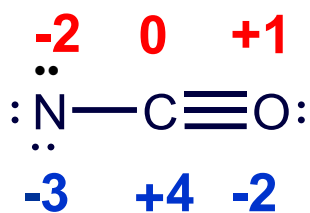
La carica formale di un atomo può essere diversa nelle diverse forme di risonanza.

$$\text{Carica formale} = \text{n.ro e}^- \text{ valenza} - (\text{n.ro e}^- \text{ di valenza non condivisi} + \frac{1}{2} \text{ n.ro e}^- \text{ valenza condivisi})$$

**Anione NCO<sup>-</sup>**

**Cariche formali**

$$\begin{array}{l} N = 5e^- \quad C = 4e^- \quad O = 6e^- \quad + 1e^- \text{ (carica)} = \frac{16e^- - 4e^-}{12e^-} \\ Si^- - (6e^- + 1e^-) = -2 \quad 6e^- - (2e^- + 2e^-) = +1 \end{array}$$



**Numeri di ossidazione**

Nella determinazione del **numero di ossidazione**, gli elettroni di legame si considerano **trasferiti** all'atomo **più** elettronegativo. Il numero di ossidazione di un atomo è lo stesso in tutte le forme di risonanza.

## Carica formale

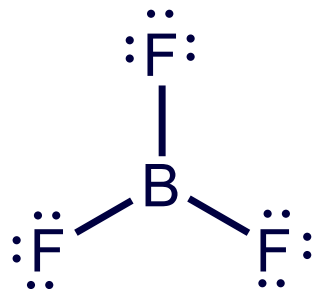
**La somma delle cariche formali (cfr definizione di numero di ossidazione) deve essere uguale alla carica effettiva della specie in questione.**

**Criteri** per stabilire le strutture di risonanza che «*meglio approssimano la verità*» o somigliano maggiormente all'ibrido:

- Le cariche formali più piccole sono preferibili a quelle più grandi (in modulo)
- Non sono desiderabili cariche simili su atomi adiacenti
- Una carica formale più negativa dovrebbe risiedere su un atomo più elettronegativo

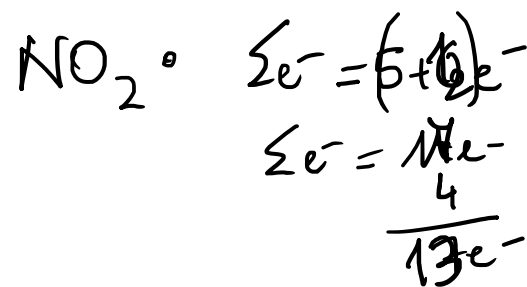
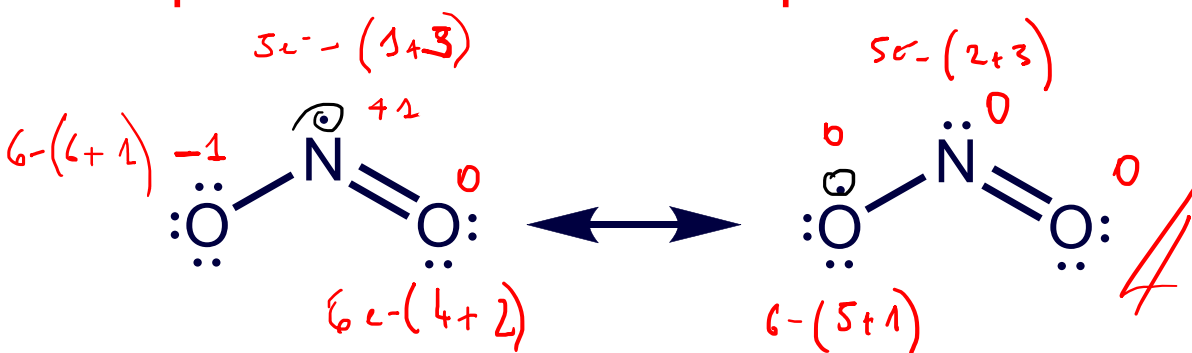
# Eccezioni alla regola dell'ottetto

## Molecole con atomi elettrone deficienti



B e Be sono elettrone deficienti.

## Specie con numero dispari di elettroni

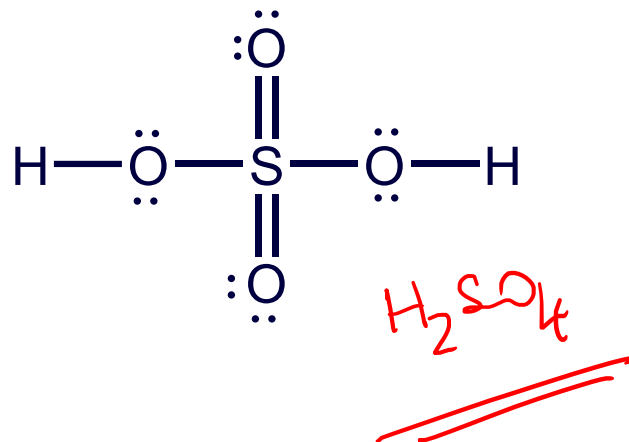
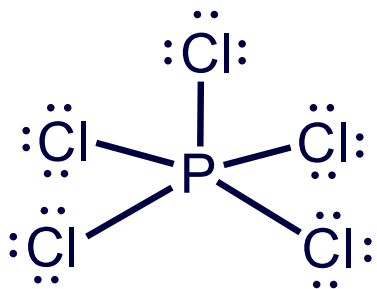


Una molecola con un numero dispari di elettroni è detta **radicale libero**.



## Eccezioni alla regola dell'ottetto

Gusci di valenza espansi



L'espansione del guscio di valenza è possibile solo per i ***non metalli del Periodo 3 o più alto*** perchè questi elementi hanno ***orbitali d*** disponibili.

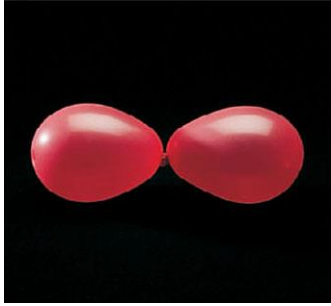
# Teoria della repulsione delle coppie di elettroni del guscio di valenza (Valence-Shell Electron-Pair Repulsion, VSEPR)

**Ciascun gruppo di elettroni di valenza attorno a un atomo centrale è situato il più lontano possibile dagli altri gruppi per minimizzare le repulsioni.**

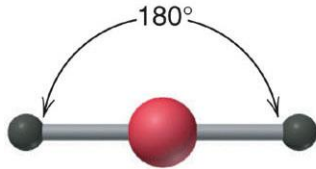
**Un “gruppo” di elettroni è qualsiasi numero di elettroni che occupano una regione localizzata attorno all'atomo.**

**Un gruppo di elettroni può essere costituito da un legame singolo, un legame doppio, un legame triplo, una coppia solitaria, o persino un elettrone solitario.**

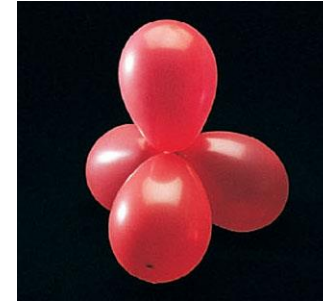
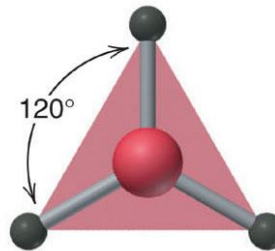
# Repulsione tra gruppi di elettroni e forme molecolari



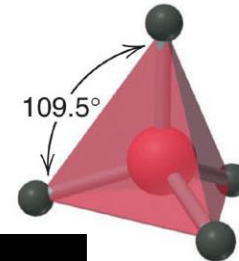
**lineare**



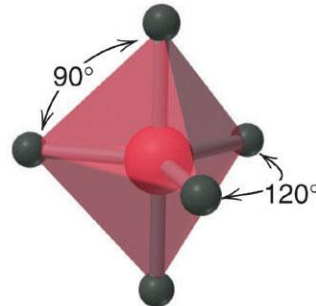
**trigonale planare**



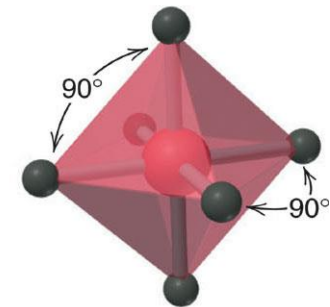
**tetraedrica**



**bipiramide  
trigonale**



**ottaedrica**



# Disposizione dei gruppi di elettroni e forma molecolare

La **disposizione dei gruppi di elettroni** è definita dai gruppi di elettroni di valenza sia di legame che di non legame.

La **forma molecolare** è la disposizione tridimensionale dei nuclei uniti dai gruppi di legame. E' definita solo dalla posizione relativa dei nuclei.

La forma molecolare può essere classificata con la notazione



A = atomo centrale

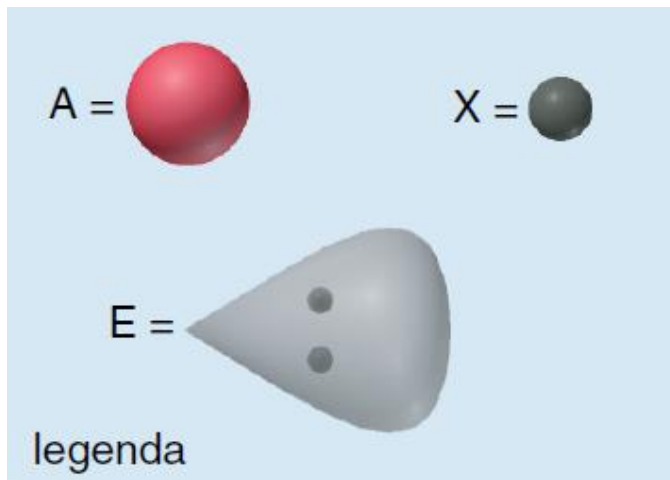
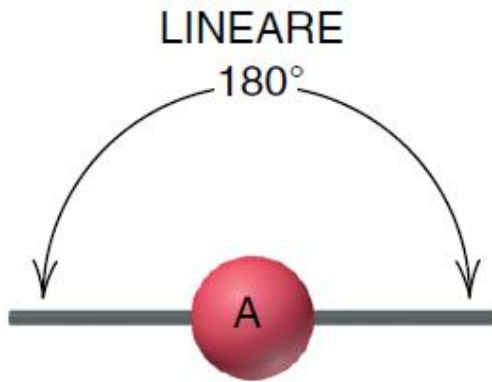
X = atomo circostante

E = gruppo di e<sup>-</sup> di valenza di non legame

*m* e *n* numeri interi

- Come determinare un modello VSEPR :
- 1. Disegnare la struttura della molecola secondo Lewis.
- 2. Contare il numero totale di coppie elettroniche intorno all'atomo centrale. Disporre le coppie in modo da minimizzare la repulsione elettrostatica.
- 3. Descrivere la geometria della molecola in termini di disposizione angolare delle coppie elettroniche di valenza.

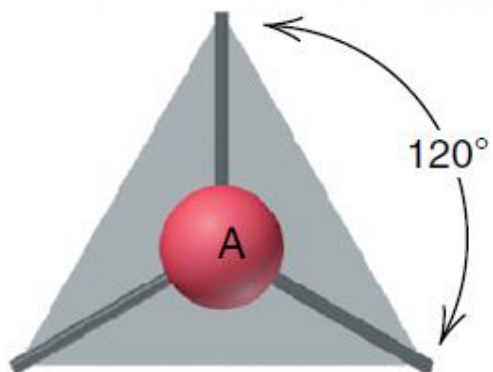
# La singola forma molecolare della disposizione lineare dei gruppi di elettroni



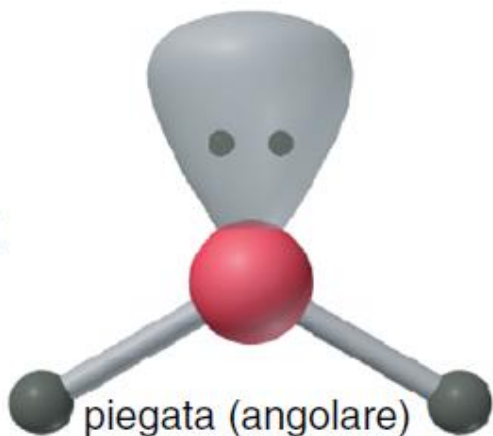
Esempi:  $CS_2$ , HCN,  $BeF_2$

# Le due forme molecolari della disposizione trigonale planare dei gruppi di elettroni

PLANARE TRIGONALE

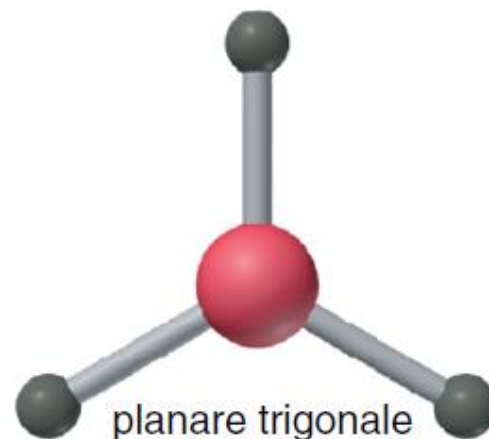


$AX_2E$



Esempi:  $SO_2$ ,  $O_3$ ,  $PbCl_2$ ,  $SnBr_2$

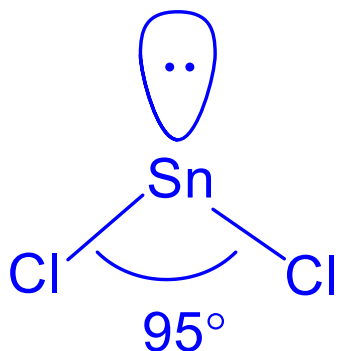
$AX_3$



Esempi:  $SO_3$ ,  $BF_3$ ,  $NO_3^-$ ,  $CO$ ,  $CO_3^{2-}$

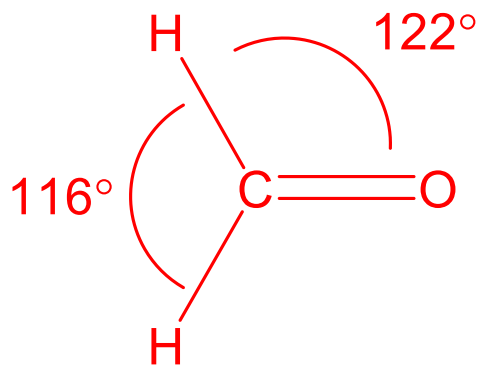
# Fattori che influenzano gli angoli di legame

## Coppie solitarie (Lone Pairs)



Una coppia solitaria respinge gli elettroni di legame più fortemente rispetto a quanto le coppie di elettroni di legame si respingano mutuamente.

## Doppi legami



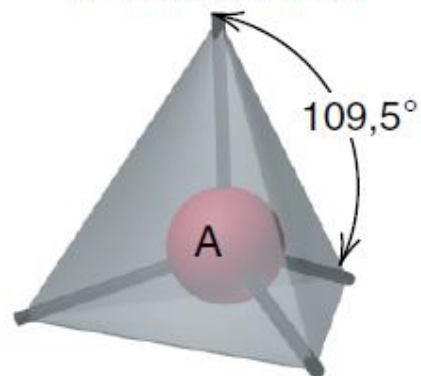
L'angolo ideale è 120°

In un doppio legame c'è una densità elettronica maggiore che in un legame singolo, la sua maggiore densità elettronica, respinge i due legami singoli più fortemente rispetto a quanto essi si respingano mutuamente.

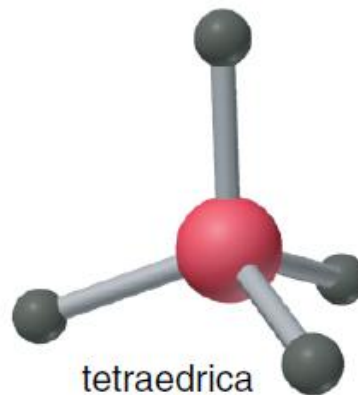


# Le tre forme molecolari della disposizione tetraedrica dei gruppi di elettroni

TETRAEDRICA

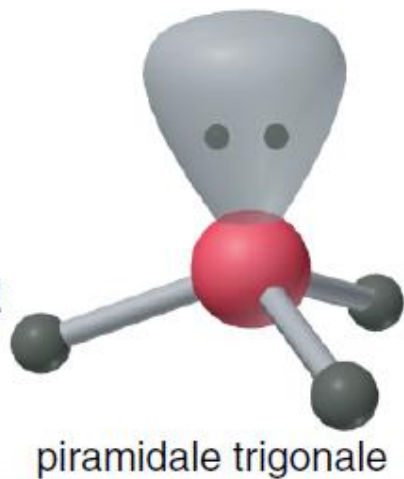


$AX_4$



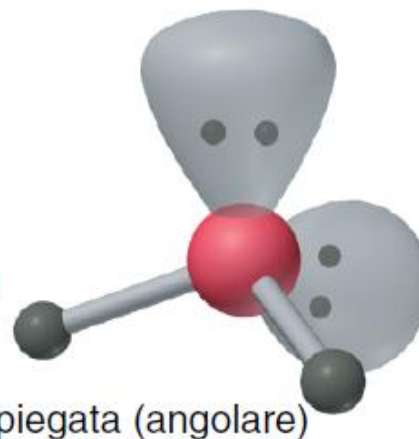
Esempi:  $CH_4$ ,  $SiCl_4$ ,  $SO_4^{2-}$ ,  $ClO_4^-$

$AX_3E$



Esempi:  $NH_3$ ,  $PF_3$ ,  $ClO_3$ ,  $H_3O^+$

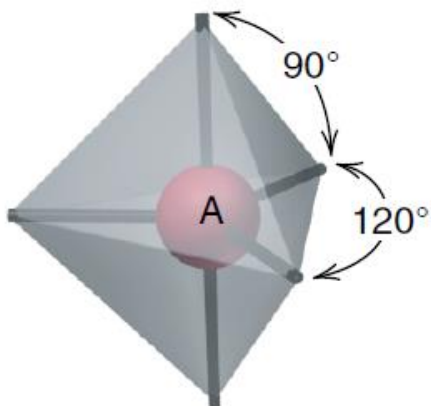
$AX_2E_2$



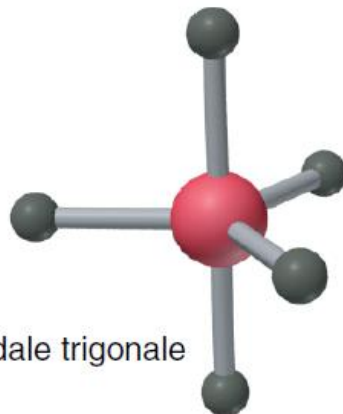
Esempi:  $H_2O$ ,  $OF_2$ ,  $SCl_2$

# Le quattro forme molecolari della disposizione a bpiramide trigonale dei gruppi di elettroni

## BIPIRAMIDALE TRIGONALE



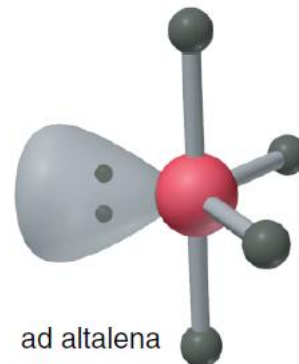
$AX_5$



bipiramidale trigonale

Esempi:  $PF_5$ ,  $AsF_5$ ,  $SOF_4$

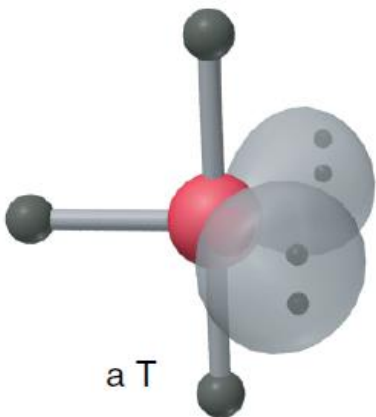
$AX_4E$



ad altalena

Esempi:  $SF_4$ ,  $XeO_2F_2$ ,  $IF_4^+$ ,  $IO_2F_2^-$

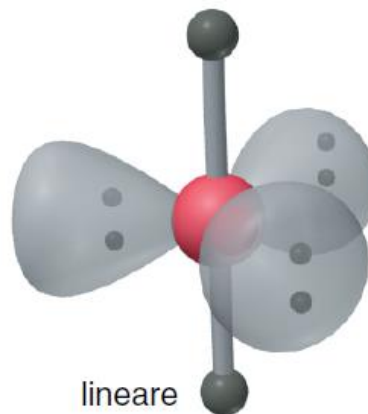
$AX_3E_2$



a T

Esempi:  $ClF_3$ ,  $BrF_3$

$AX_2E_3$

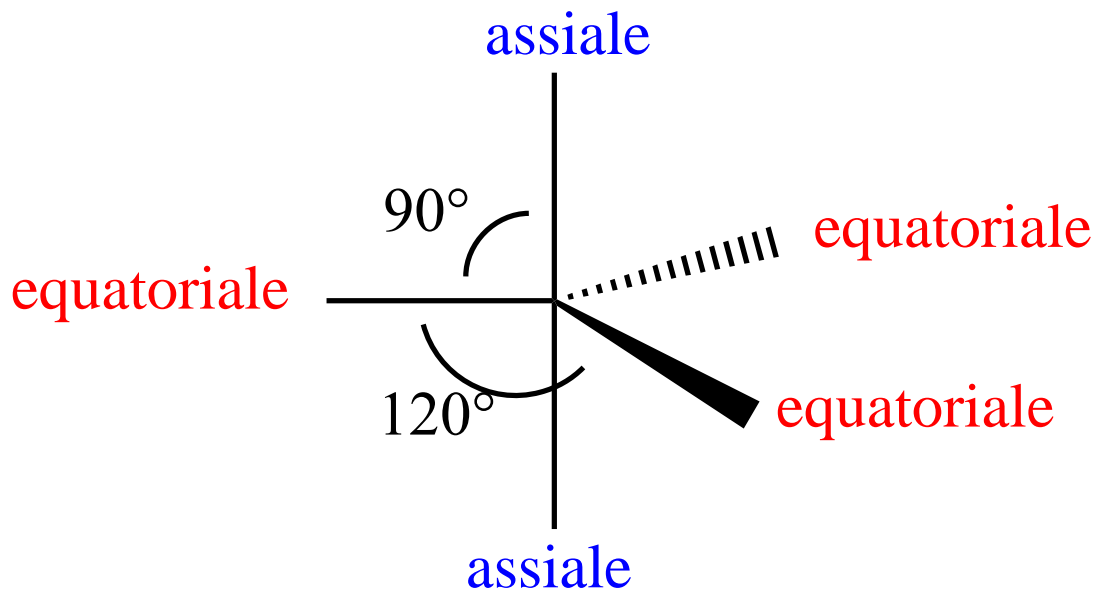


lineare

Esempi:  $XeF_2$ ,  $I_3^-$ ,  $IF_2^-$

## Posizioni assiali ed equatoriali

Un sistema con 5 gruppi di elettroni ha due diverse posizioni per i gruppi di elettroni e due angoli ideali

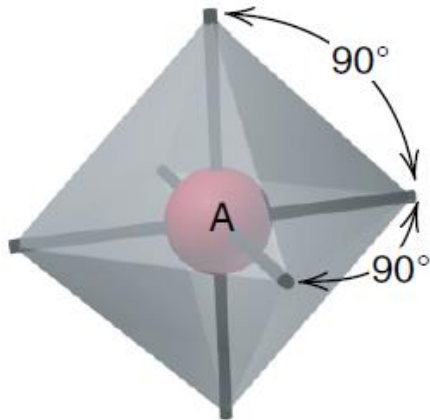


Le repulsioni equatoriale-equatoriale sono più deboli delle repulsioni equatoriale-assiale.

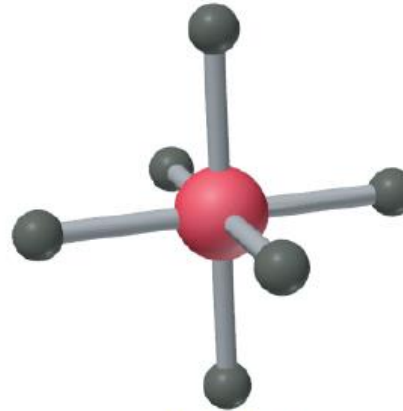
Quando possibile, i lone pair in un sistema a cinque gruppi di elettroni occupano le posizioni **equatoriali**.

# Le tre forme molecolari della disposizione ottaedrica dei gruppi di elettroni

OTTAEDRICA



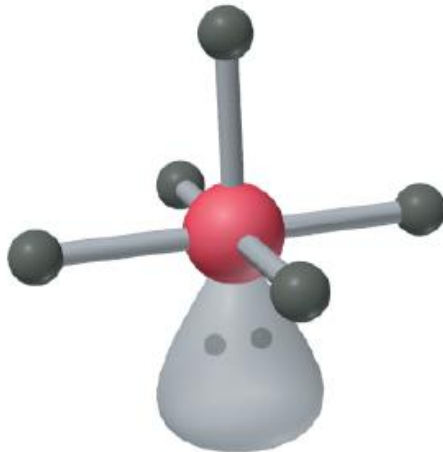
$AX_6$



ottaedrica

Esempi:  $SF_6$ ,  $IOF_5$

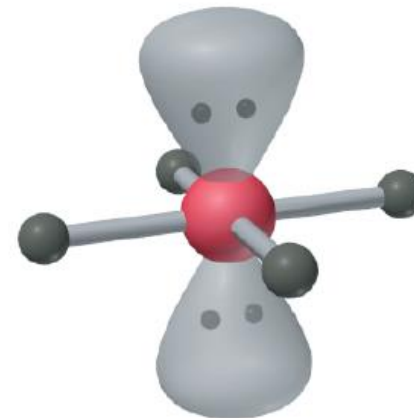
$AX_5E$



piramidale quadrata

Esempi:  $BrF_5$ ,  $TeF_5^-$ ,  $XeOF_4$

$AX_4E_2$



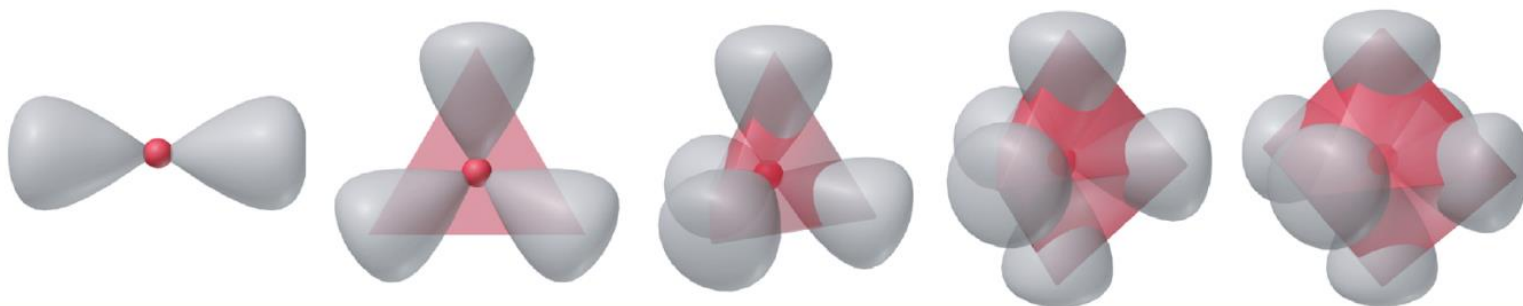
planare quadrata

Esempi:  $XeF_4$ ,  $ICl_4^-$

# Composizione e orientazione degli orbitali ibridi

|                                       | Lineare                      | Planare trigonale            | Tetraedrica                    | Bipiramidale trigonale                       | Ottaedrica                                       |
|---------------------------------------|------------------------------|------------------------------|--------------------------------|--|--|
| <b>Orbitali atomici mescolati</b>     | uno <i>s</i><br>uno <i>p</i> | uno <i>s</i><br>due <i>p</i> | uno <i>s</i><br>tre <i>p</i>   | uno <i>s</i><br>tre <i>p</i><br>uno <i>d</i> | uno <i>s</i><br>tre <i>p</i><br>due <i>d</i>     |
| <b>Orbitali ibridi formati</b>        | due <i>sp</i>                | tre <i>sp</i> <sup>2</sup>   | quattro <i>sp</i> <sup>3</sup> | cinque <i>sp</i> <sup>3</sup> <i>d</i>       | sei <i>sp</i> <sup>3</sup> <i>d</i> <sup>2</sup> |
| <b>Orbitali non ibridati restanti</b> | due <i>p</i>                 | uno <i>p</i>                 | nessuno                        | quattro <i>d</i>                             | tre <i>d</i>                                     |

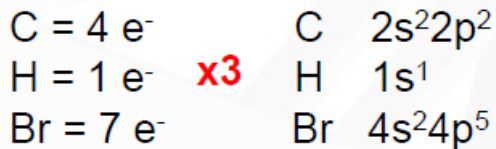
Orientamento



Dalla VSEPR posso ottenere info sugli **orbitali ibridi** utilizzati dall'atomo centrale

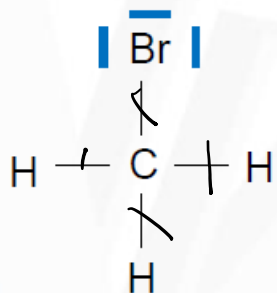
# ESEMPI

CH<sub>3</sub>Br



**14 e<sup>-</sup>** → 7 doppietti

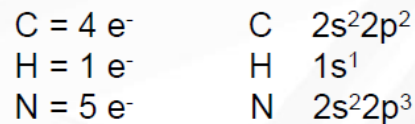
In questo caso **C** è l'atomo centrale



4 doppietti usati → **8e<sup>-</sup> usati**

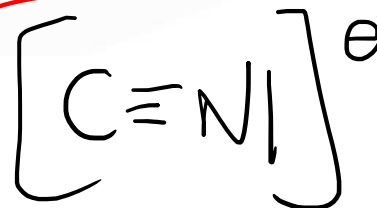
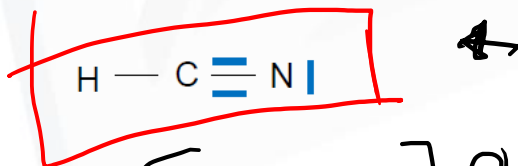
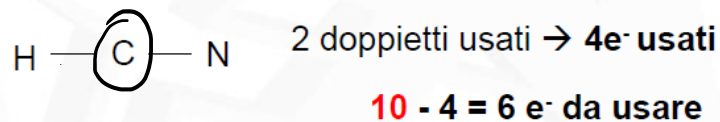
**14 - 8 = 6 e<sup>-</sup> da usare**

HCN



**10 e<sup>-</sup>** → 5 doppietti

In questo caso **C** è l'atomo centrale



# ESEMPI



a. Scrivere formula Lewis

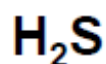


---

$$48 e^- \rightarrow 24 \text{ doppietti}$$

6 doppietti usati  $\rightarrow$  **12e<sup>-</sup> usati**

$$48 - 12 = 36 e^- \text{ da usare}$$



a. Scrivere formula Lewis

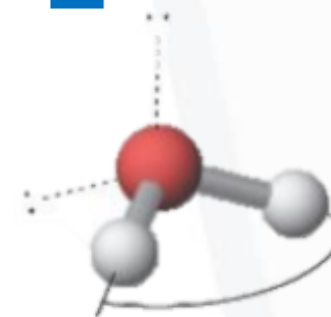
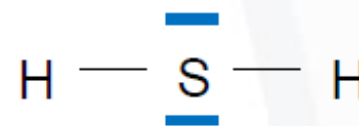
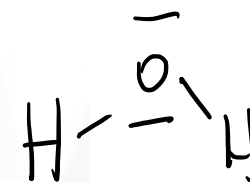
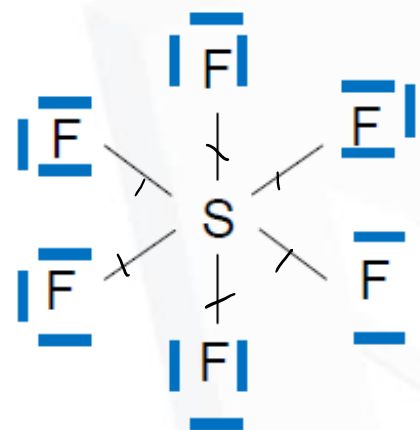


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$$8 e^- \rightarrow 4 \text{ doppietti}$$

2 doppietti usati  $\rightarrow$  **4 e<sup>-</sup> usati**

$$8 - 4 = 4 e^- \text{ da usare}$$



$\text{HSO}_4^-$  come negativo

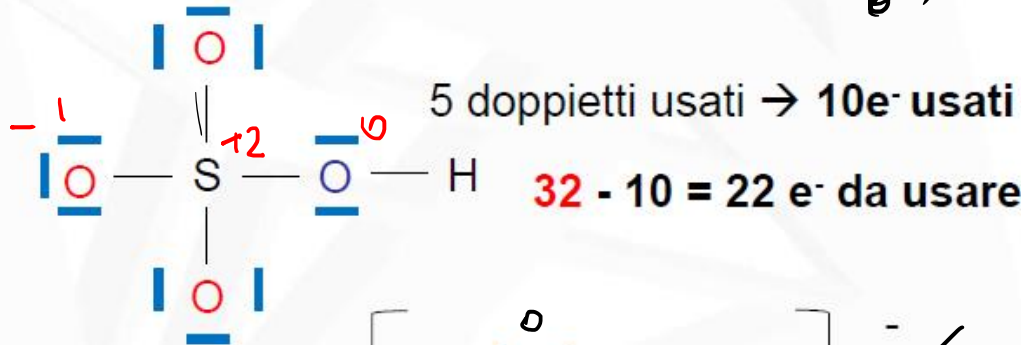
$(\text{H}_3\text{PO}_3)$

H = 1 e<sup>-</sup>      H 1s<sup>1</sup>  
 S = 6 e<sup>-</sup>      S 3s<sup>2</sup>3p<sup>4</sup>  
 O = 6 e<sup>-</sup> x4    O 2s<sup>2</sup>2p<sup>4</sup>

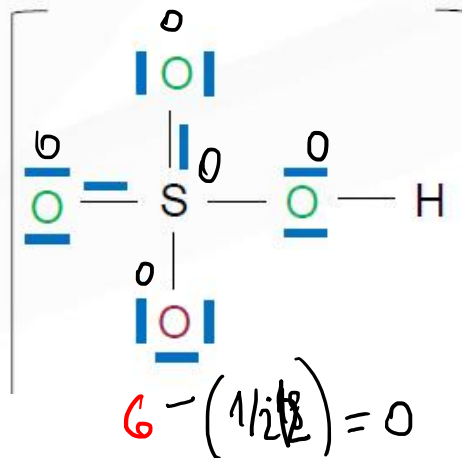
+ 1 e<sup>-</sup> (carica negativa)

32 e<sup>-</sup> → 16 doppietti

In questo caso S è l'atomo centrale



Carica formale  
 S = 0  
 3 O = 0  
 1 O = -1 (carica -)



6/4



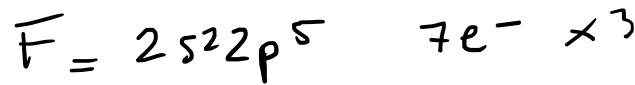
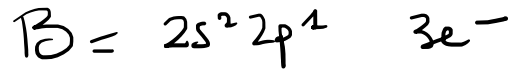
$\text{H}_x\text{N}_y\text{O}_z$

H acidi è sempre legato all'O

- HCl
- HBr
- HI
- H<sub>2</sub>S

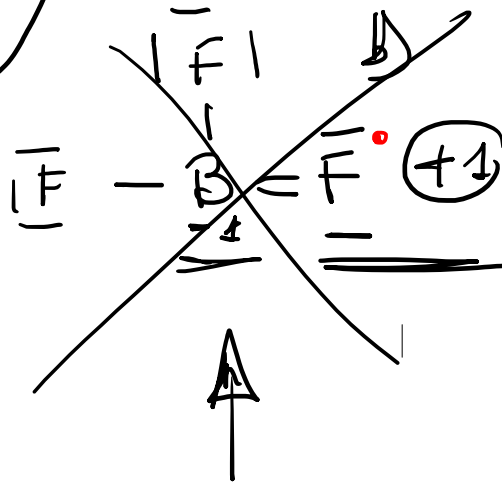
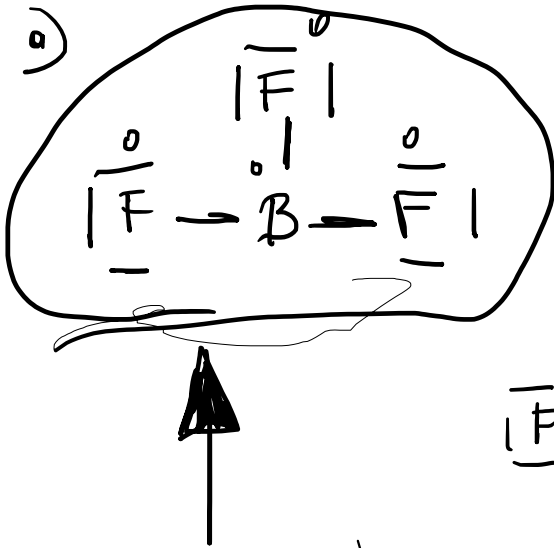
$6e^- - (0 + 1/2 \cdot 8) = +2$   
 $6e^- - (6 + 1/2 \cdot 2) = -1$   
 $6e^- - (4 + 1/2 \cdot 2) = 0$   
 $6e^- - (4 + 1/2 \cdot 4) = 0$





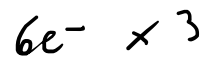
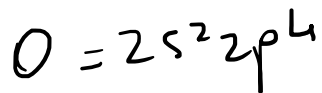
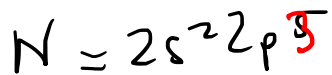
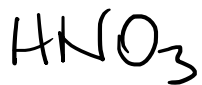
$$\Sigma e^{-} = \frac{24e^{-}}{6e^{-}} = \frac{18e^{-}}{18e^{-}}$$

Regola ottetto

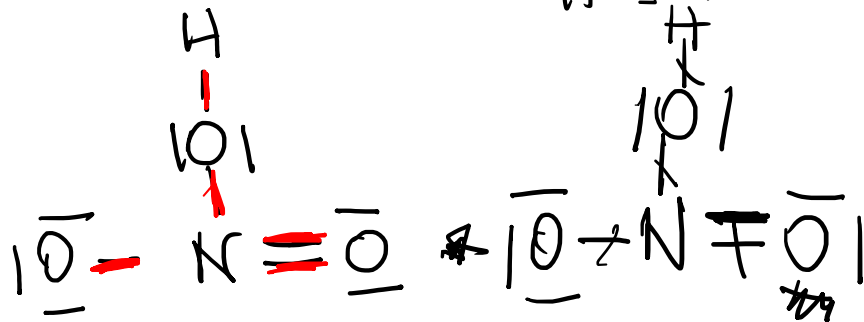


a) B = 3e<sup>-</sup> - (1/2 · 6) = 0  
 F = 7e<sup>-</sup> - (6 + 1/2 · 2) = 0

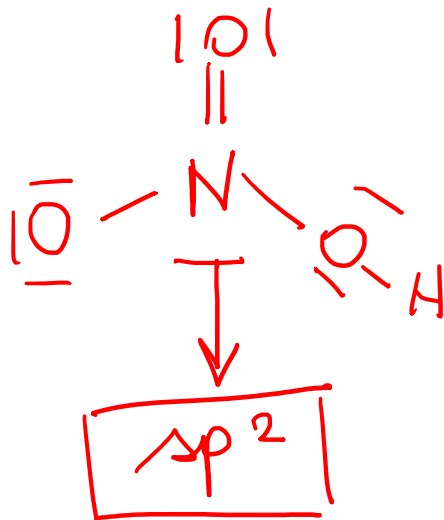
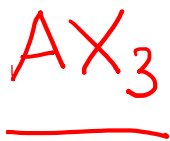
b) B = 3e<sup>-</sup> - (1/2 · 8) = -1  
 F = 0  
 F = 7 - (4 + 1/2 · 4) = +1



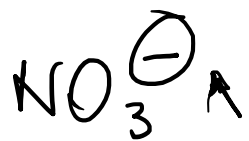
$$\begin{array}{r} \hline \Sigma_{\text{tot}} = \frac{24e^-}{8e^-} \\ \hline \underline{\underline{16e^-}} \end{array}$$



⇓ Geometria



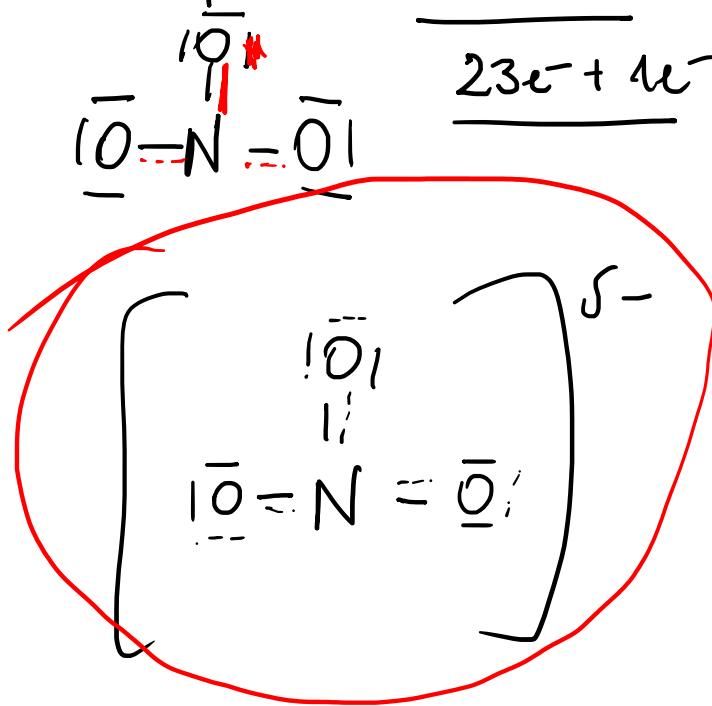
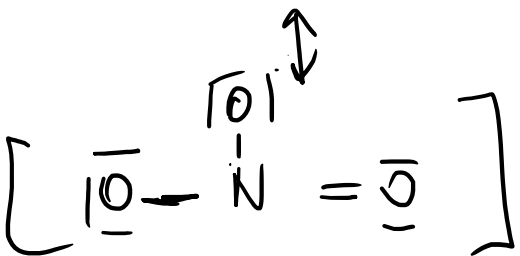
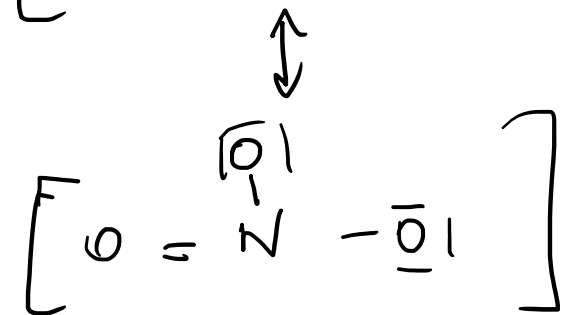
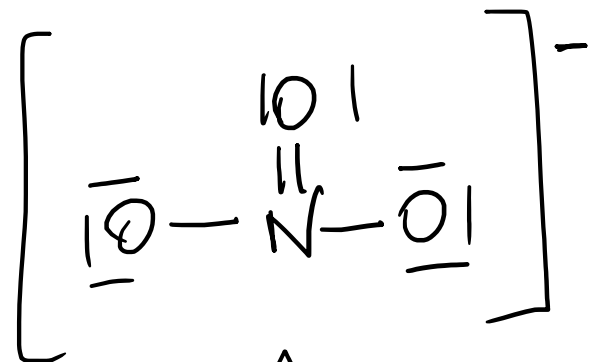
TRIGONALE  
PIANALE



$$N = 2s^2 2p^3 = 5e^-$$

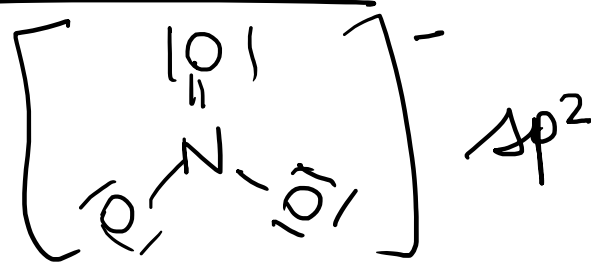
$$O = 2s^2 2p^4 = 6e^- \times 3$$

$$\frac{23e^- + 4e^- = 27e^- - 6e^-}{18e^-}$$



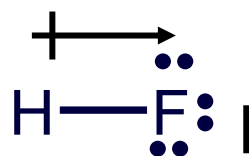
IBRIDO di  
RISONANZA

$AX_3$

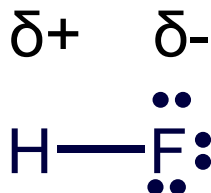


## Rappresentazione dei legami polari

La condivisione ineguale degli elettroni può essere rappresentata con una freccia polare orientata verso *l'elemento più elettronegativo*.



Un legame polare può anche essere rappresentato con i simboli  $\delta+$  e  $\delta-$  che indicano una carica parziale.



# Momento di dipolo elettrico, $\mu$

Indica l'entità dello spostamento della carica in un legame covalente polare:

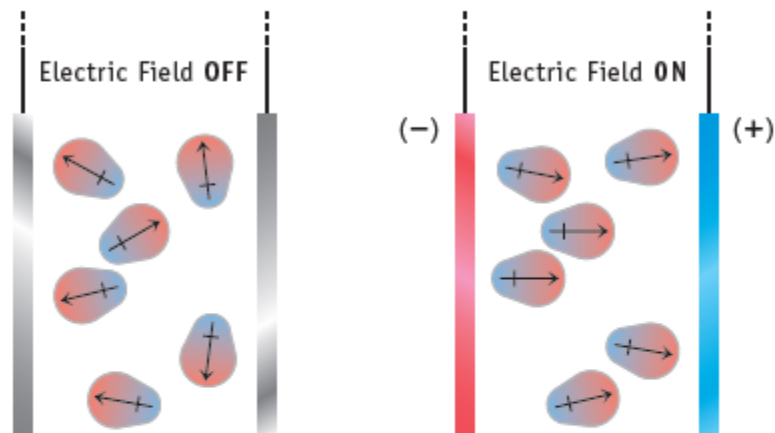
$$\mu = \delta \times d$$

## A CLOSER LOOK

## Polar Molecules in an Electric Field

Let us look at the experimental measurement of the polarity of a molecule. When placed in an electric field created by a pair of oppositely charged plates, polar molecules experience a force that tends to align them with the field. The positive end of each molecule is attracted to the negative plate, and the negative end is attracted to the positive plate. The orientation of the polar molecule affects the electrical capacitance of the plates (their ability to hold a charge), and this provides a way to measure experimentally the magnitude of the dipole moment.

The extent to which the molecules line up with the field depends on their **dipole moment**,  $\mu$ . In a molecule with partial charges equal to  $+q$  and  $-q$  and in which these charges are separated by a distance,  $d$ , in the molecule, the magnitude of the dipole moment is given by the equation  $\mu = q \times d$ . The SI unit of the dipole moment is the



Polar molecules in an electric field.

coulomb-meter, but dipole moments have traditionally been given using a derived unit called the *debye* (D;  $1 \text{ D} = 3.34 \times 10^{-30} \text{ C} \cdot \text{m}$ ). Experimental values of some dipole moments are listed in Table 8.6.

The commonly used unit of dipole moments is named in honor of Peter Debye

(1884–1966). Debye was born in the Netherlands and educated in Germany. He received the Nobel Prize in Chemistry in 1936 and was later a professor at Cornell University.

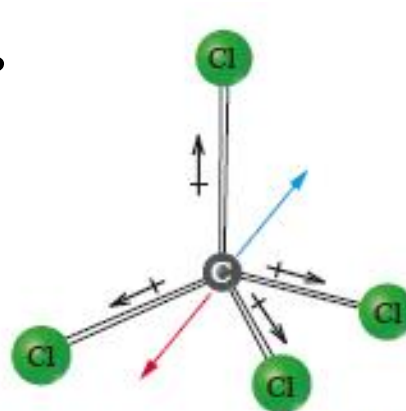
# VSEPR e polarità delle molecole

*Non sempre la presenza di legami covalenti polari dà origine ad una molecola polare*

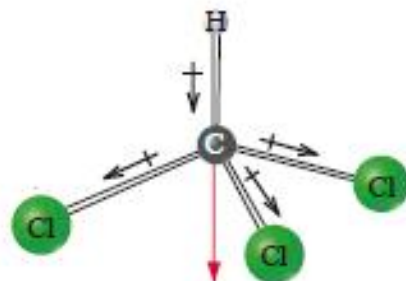
► FIGURA 10-15

## Forme molecolari e momenti di dipolo

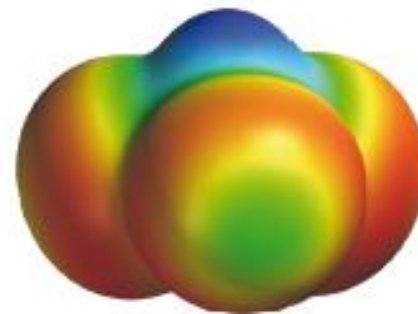
(a) La risultante di due dei dipoli C—Cl è indicata dalla freccia rossa e quella degli altri due da quella blu. Le frecce rossa e blu puntano in direzioni opposte e si annullano. La molecola  $\text{CCl}_4$  è apolare. Il bilanciamento della distribuzione di carica in  $\text{CCl}_4$  si vede bene nella mappa del potenziale elettrostatico. (b) I singoli dipoli individuali si combinano dando un momento di dipolo risultante (freccia rossa) di 1.04 D. La mappa del potenziale elettrostatico indica che l'atomo di idrogeno ha una carica parziale positiva.

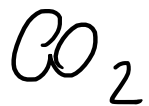


(a)  $\text{CCl}_4$ : una molecola apolare



(b)  $\text{CHCl}_3$ : una molecola polare

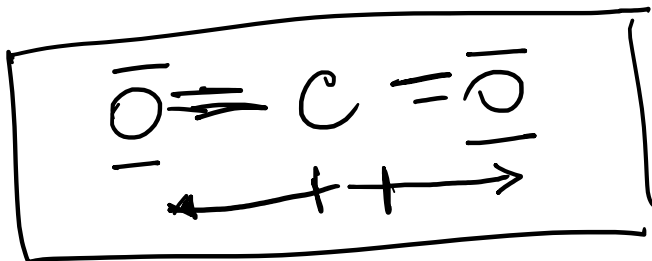




$$C = 4e^-$$

$$O = 6e^- \times 2 = 12e^-$$

$$16e^- - 4e^- = 12e^-$$

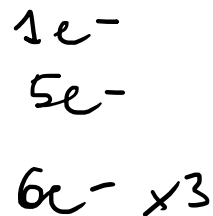
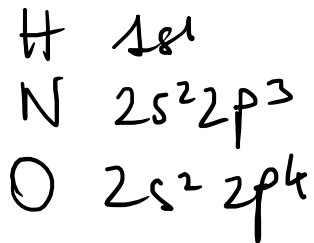


AX<sub>2</sub> planare

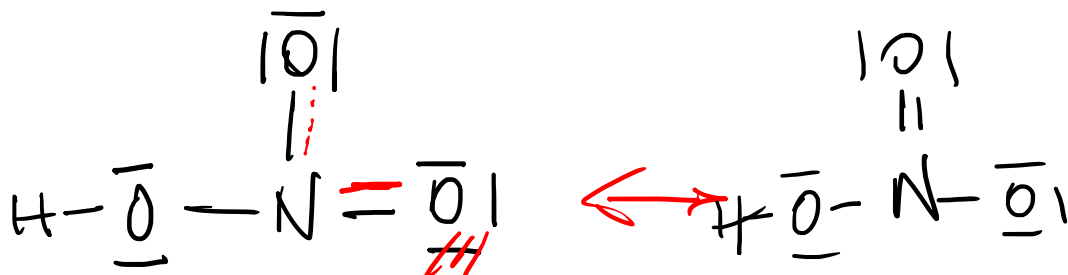
Apolare

SCARSA REATTIVITA'

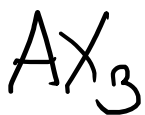
EX VSEPR



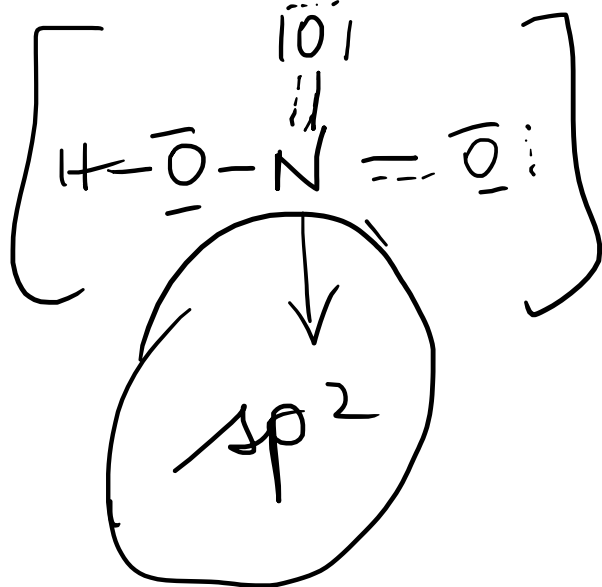
$$\begin{array}{r} \Sigma e^- = 24 e^- \\ \underline{8e^-} \\ 16e^- \end{array}$$



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RISONANZA

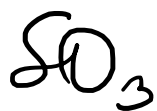


Trigonale  
planare  
( $120^\circ\text{C}$ )



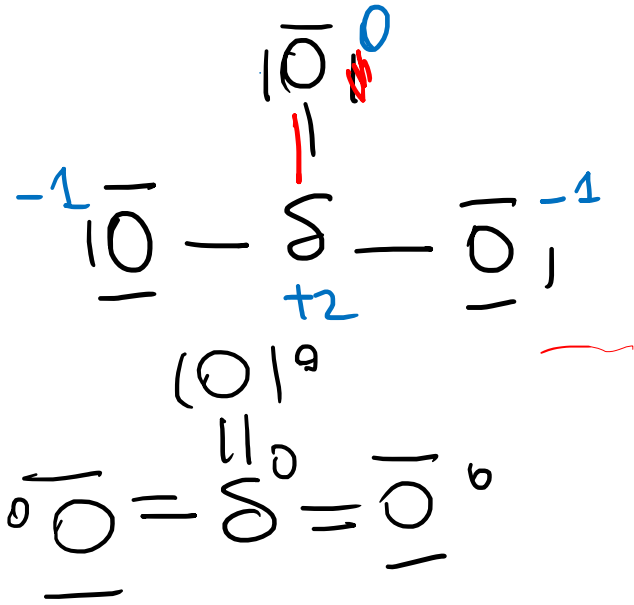
①





$$\begin{aligned} \Sigma & \textcircled{3s^2} \textcircled{3p^4} = 6e^- \\ 0 & 2s^2 2p^4 = 6e^- \times 3 \end{aligned}$$

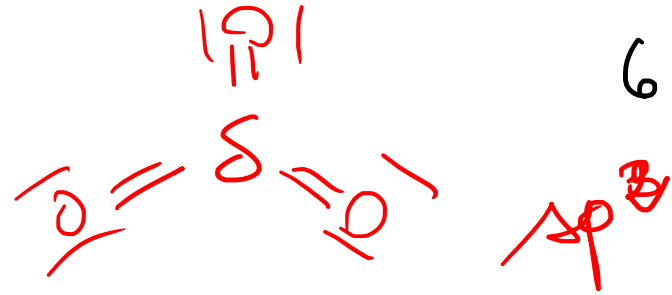
$$\begin{array}{r} \Sigma e^- \\ \hline 24e^- \\ \hline 6e^- \\ \hline 18e^- \end{array}$$



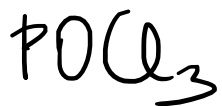
$$\begin{aligned} cf_s &= 6 - (0 + 1/2 \cdot 6) = \textcircled{+2} \\ cf_o &= 6 - (6 + 1/2 \cdot 2) = -1 \\ cf_{ob} &= 6 - (4 + 1/2 \cdot 4) = 0 \end{aligned}$$

$$6 - (0 + 1/2 \cdot 12) = 0$$

AX3



$\textcircled{d}$



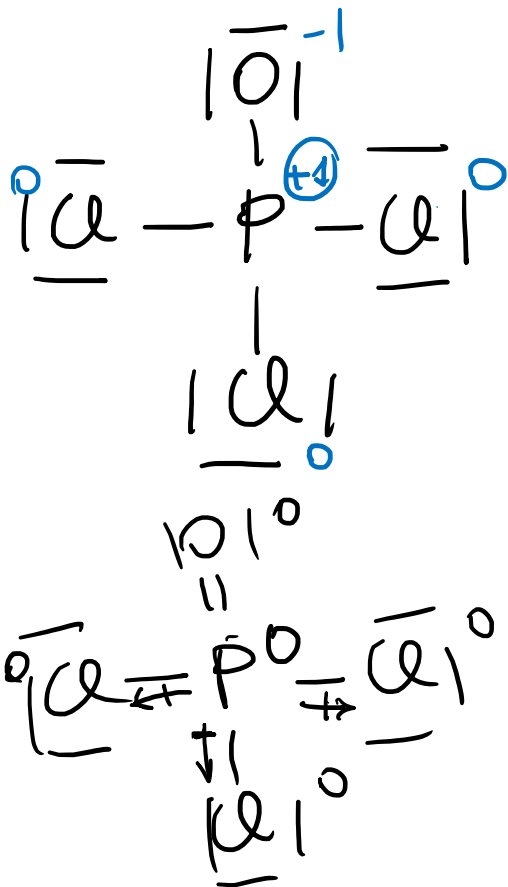
$\text{P } 3s^2 3p^3 = 5e^-$



$\text{O } 2s^2 2p^4 = 6e^-$

$\text{Cl } 3s^2 3p^5 = \frac{7e^- \times 3}{5e^- \quad 32e^-}$

$\frac{7e^-}{24e^-}$



POLARE

$\text{P} = 5e^- - (0 + 1/2 \cdot 8) = +1$   
 $\text{Cl} = 7e^- - (6 + 1/2 \cdot 2) = 0$   
 $\text{O} = 6e^- - (6 + 1/2 \cdot 2) = -1$

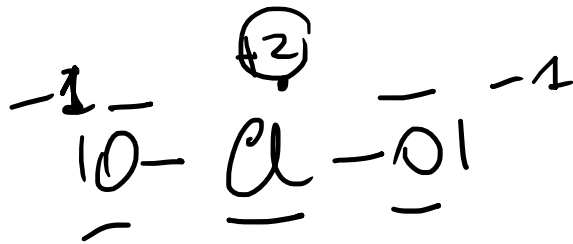
$sp^3$  AX<sub>4</sub> (d)



C 3s<sup>2</sup> 3p<sup>5</sup> = 7 e<sup>-</sup>

O 2s<sup>2</sup> 2p<sup>4</sup> = 6e<sup>-</sup> x 2

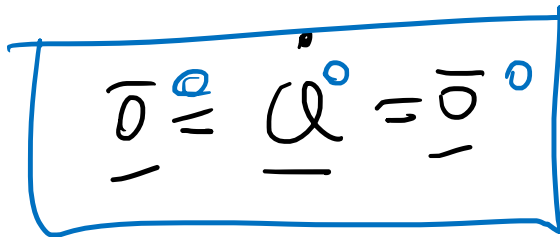
$$\begin{array}{r} 2e^- \quad 18e^- \\ \hline 4e^- \\ \hline 18e^- \end{array}$$



AX<sub>2</sub>E<sub>2</sub>

C = 7 - (3 + 1/2 \* 4) = +2

O = 6 - (6 + 1/2 \* 2) = -1

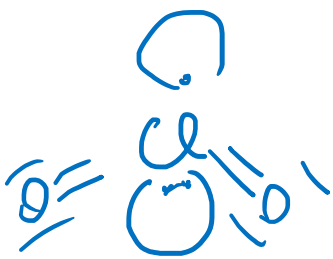


C = 7 - (3 + 1/2 \* 8) = 0

O = 6 - (4 + 1/2 \* 4) = 0

sp<sup>3</sup>

(d)



pregata  
anzilor

↑  
Formula vera

CO

$$C 2s^2 2p^2 = 4e^-$$

$$O 2s^2 2p^4 = 6e^-$$

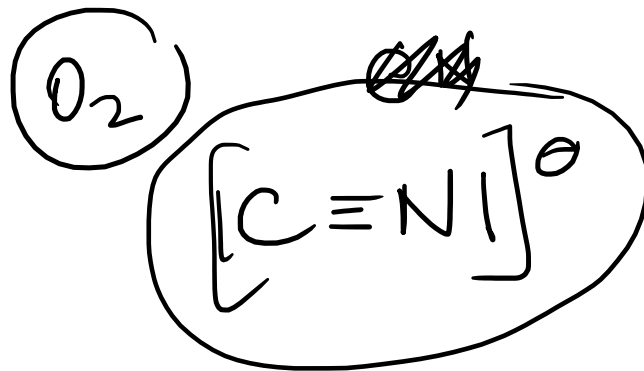
$$\frac{\overline{\Sigma 10e^-}}{2e^-} = 5e^-$$

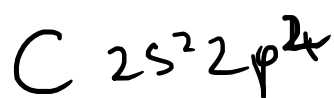
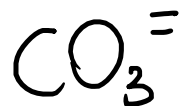
$$|\bar{C} - \bar{O}|$$

$$|\bar{C} = \bar{O}|$$

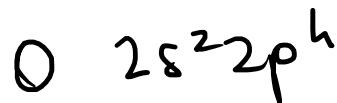
$$|C \equiv O|$$

sp (p)





$$= 4e^-$$



$$= \frac{6e^- \times 3}{}$$

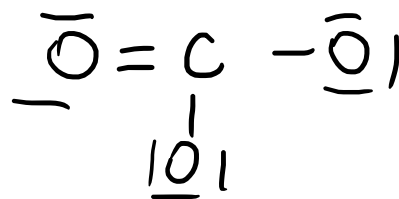
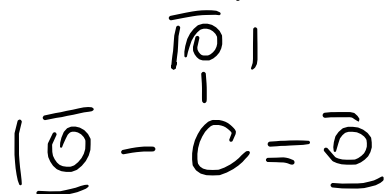
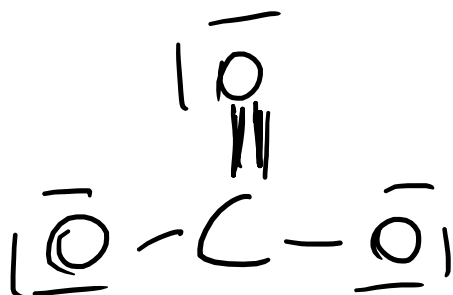
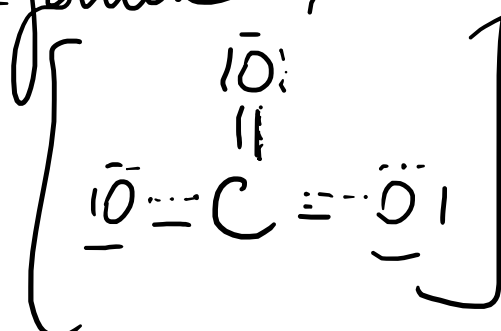
$$\Sigma e^- = 22e^- +$$

$$2e^-$$

$$\boxed{24e^-}$$

$$\frac{6e^-}{18e^-}$$

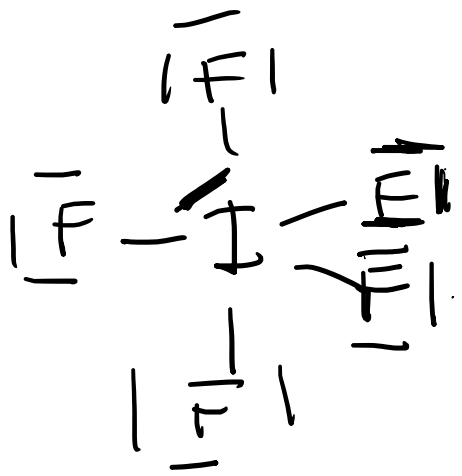
3 famille d'isomérie



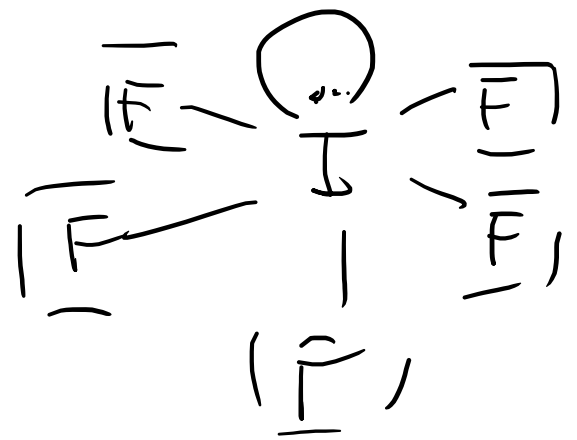
IF<sub>5</sub>

$$\begin{array}{r} \text{I} \quad 5s^2 5p^5 = 7e^- \\ \text{F} \quad 2s^2 2p^5 = 7e^- \times 5 \\ \hline \end{array}$$

$$\begin{array}{r} 42e^- \\ 10e^- \\ \hline 32e^- \\ \hline \end{array}$$



AX<sub>5</sub>E



sp<sup>3</sup>d<sup>2</sup>