

Bonding & Molecular Geometry

CHEM 50 · Dr. Karmach

By the end of this unit, you can...

- Classify a bond as ionic, polar covalent, or nonpolar covalent from the electronegativity difference
- Write the Lewis (electron dot) symbol of an atom or a monatomic ion
- Count valence electrons and draw valid Lewis structures for molecules and polyatomic ions
- Draw all resonance structures of a molecule or ion and describe the blend they represent
- Assign formal charges and use them to choose the best Lewis structure
- Draw the octet-rule exceptions: incomplete octets, odd-electron molecules, and expanded octets
- Apply VSEPR theory to predict electron-pair geometry and molecular shape
- Combine bond dipoles with molecular shape to judge whether a molecule is polar
- Assign trigonal bipyramidal and octahedral electron geometries and their lone-pair shapes, and judge the polarity of five- and six-domain molecules

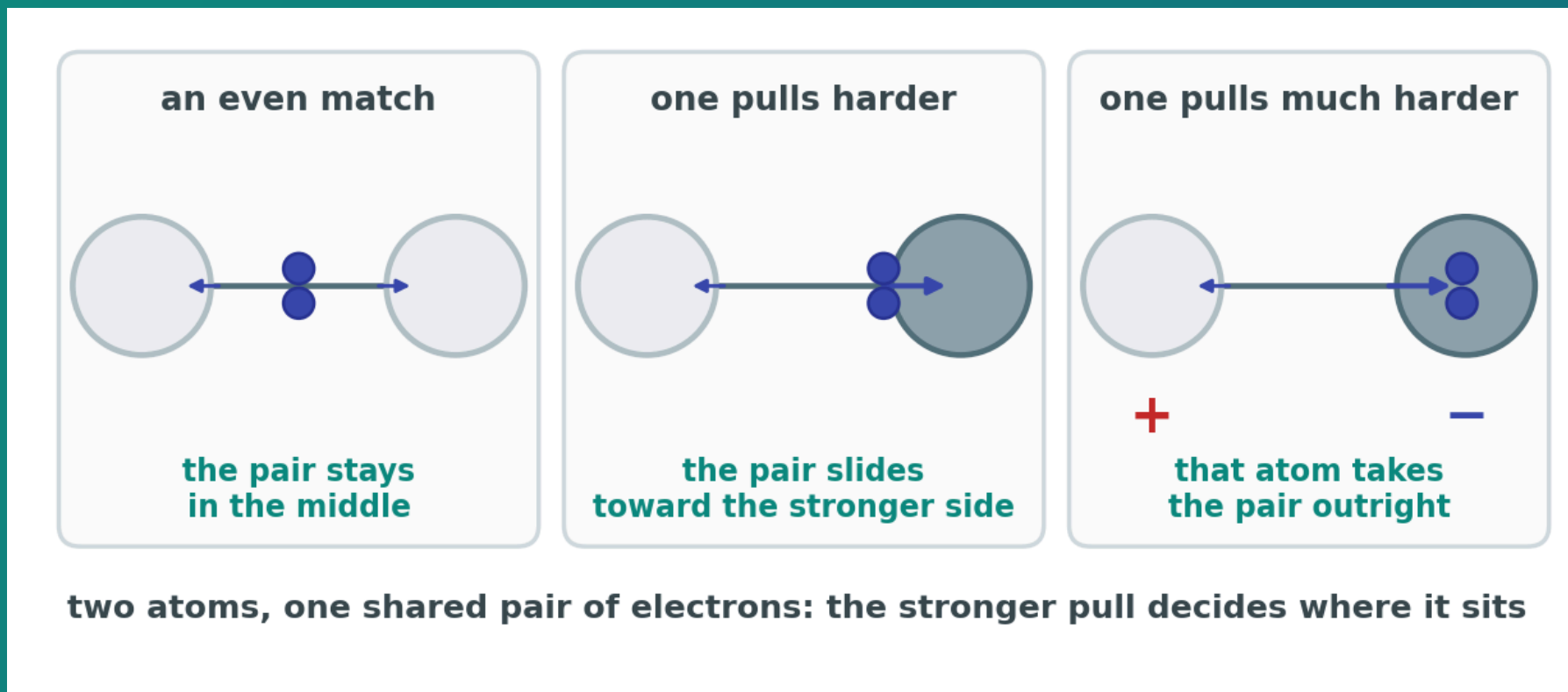
Today's route

1. **Electronegativity & Bond Type**
2. **Lewis Structures**
3. **Resonance & Formal Charge**
4. **Exceptions to the Octet Rule**
5. **VSEPR & Molecular Shape**
6. **Molecular Polarity**
7. **Five & Six Electron Domains**
8. **Lattice Energy Trends**
9. **Bond Energies & Reaction Enthalpy**

1 • Electronegativity & Bond Type

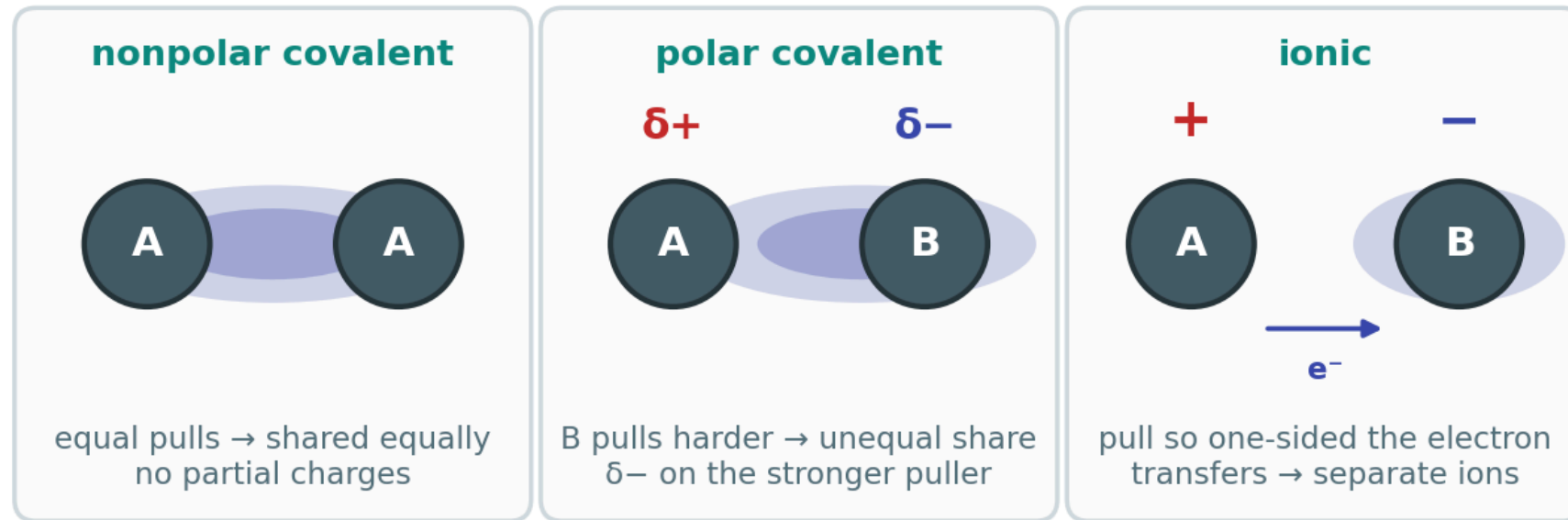
Classify any bond as nonpolar covalent, polar covalent, or ionic from the electronegativity difference, and mark which atom carries the partial negative charge.

Pulling on shared electrons



Two atoms pull on the same pair of electrons. Pull evenly, and it stays centered. Pull much harder, and one atom drags the pair over, or takes it.

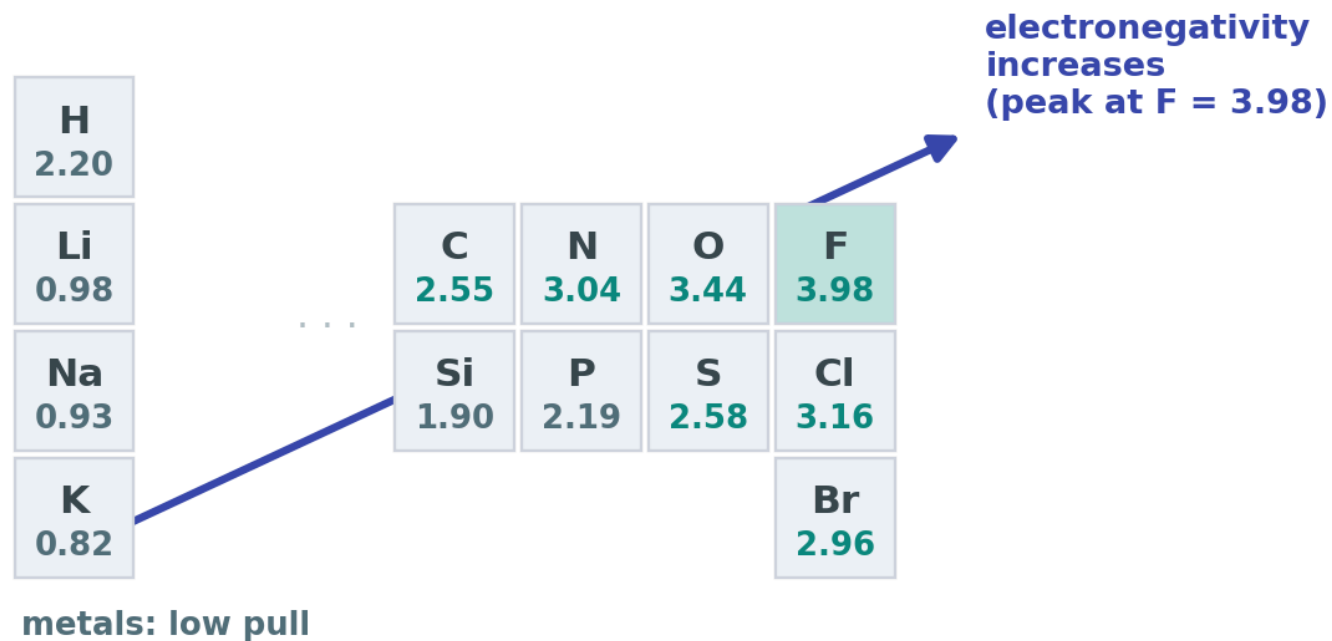
How a bond shares its electrons



a bond is one shared pair: how equally the atoms share it names the bond

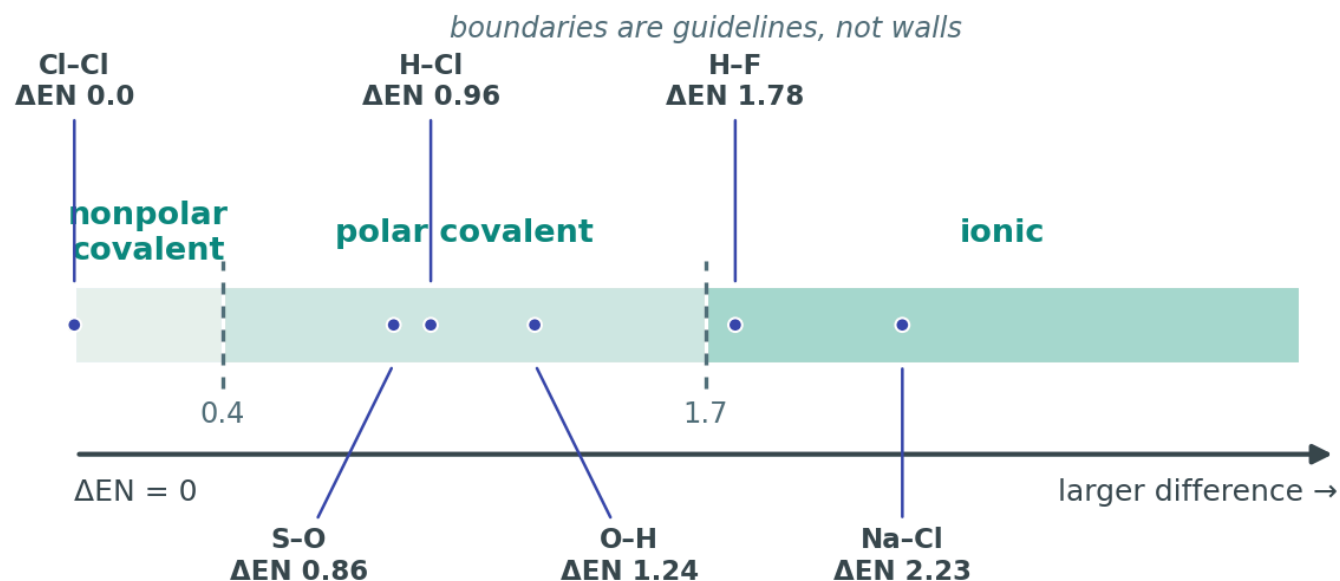
A bond is a shared pair of electrons. Each atom pulls on that pair; the strength of the pull is its electronegativity. Equal pulls share equally; unequal pulls share unequally.

Electronegativity and its trend



Electronegativity measures how strongly an atom pulls on shared electrons. It climbs across a row and up a column, peaking at fluorine. Metals sit low; nonmetals near fluorine sit high.

From difference to bond type



Subtract the two electronegativities, larger minus smaller, and read the bond type off the continuum.

under 0.4: even sharing • 0.4 to 1.7: unequal sharing • past 1.7: transfer
nonpolar covalent • polar covalent • ionic • guidelines, not walls



The method

1. **Find the electronegativity difference.** Larger minus smaller.
2. **Place it on the continuum.** Near zero, nonpolar covalent; moderate, polar covalent; large, ionic.
3. **Name the bond and mark the charges.** δ^- on the more electronegative atom, δ^+ on its partner.



Worked example 1: the Cl-Cl bond

Cl-Cl (the bond inside Cl₂ gas)

electronegativity: Cl = 3.16 · Cl = 3.16 · wanted: bond type

Chlorine gas is two chlorine atoms sharing one pair of electrons.

Classify the bond.

✓ Worked example 1: solution

Cl-Cl (the bond inside Cl₂ gas)
electronegativity: Cl = 3.16 · Cl = 3.16

Step 1 · Find the electronegativity difference

$$\Delta\text{EN} = 3.16 - 3.16 = 0$$

Step 2 · Place it on the continuum

Step 3 · Name the bond and mark the charges

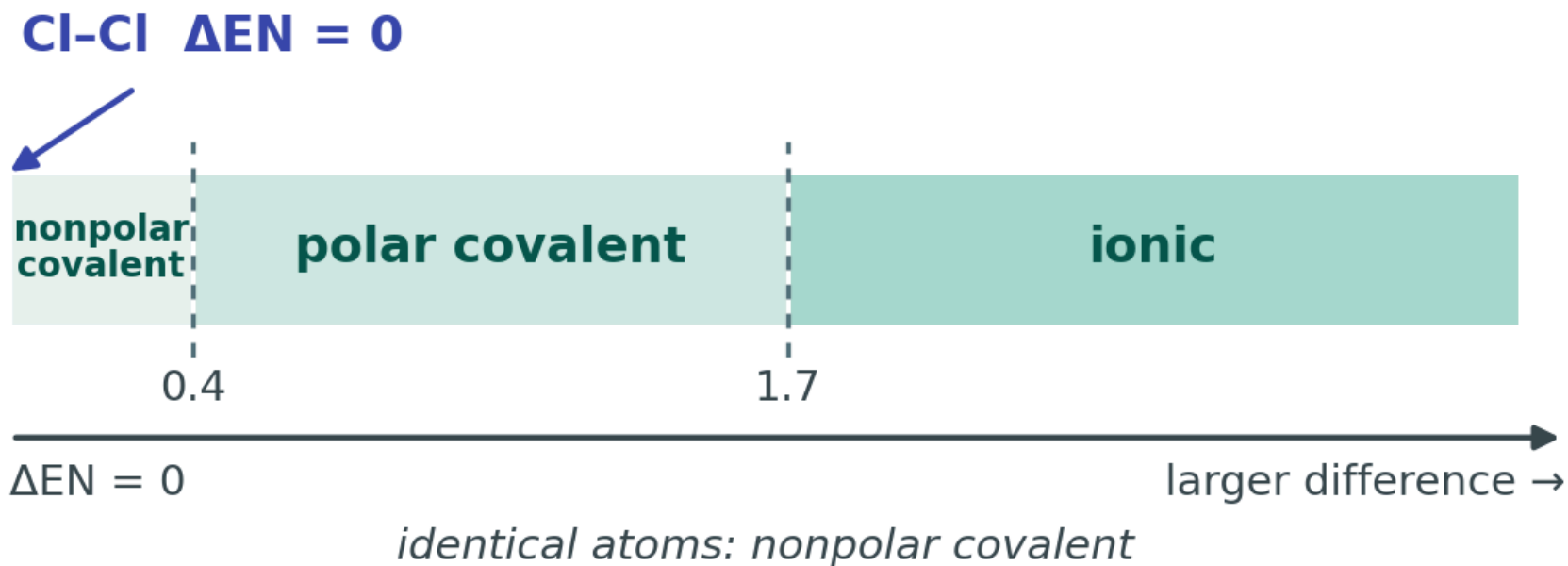
Cl-Cl → nonpolar covalent
equal electronegativities · the pair sits centered · no partial charges

 **Sense check:** Two atoms of the same element always pull equally. A bond between identical atoms is nonpolar, every time.

✓ Worked example 1: the bond on the continuum

$$\text{Cl-Cl: } \Delta\text{EN} = 3.16 - 3.16 = 0$$

electronegativity: Cl = 3.16 · Cl = 3.16 · found: nonpolar covalent



 **Sense check:** Zero sits at the far left end of the continuum. Identical atoms share the pair evenly.



Worked example 2: the H-Cl bond

H-Cl (hydrogen chloride)

electronegativity: H = 2.20 · Cl = 3.16 · wanted: bond type and partial charges

Hydrogen and chlorine share one pair of electrons.

A common first answer marks hydrogen as δ^- . Test it against the two electronegativities.

✓ Worked example 2: solution

H-Cl (hydrogen chloride)
electronegativity: H = 2.20 · Cl = 3.16

Step 1 · Find the electronegativity difference

$$\Delta\text{EN} = 3.16 - 2.20 = 0.96$$

Step 2 · Place it on the continuum

Step 3 · Name the bond and mark the charges

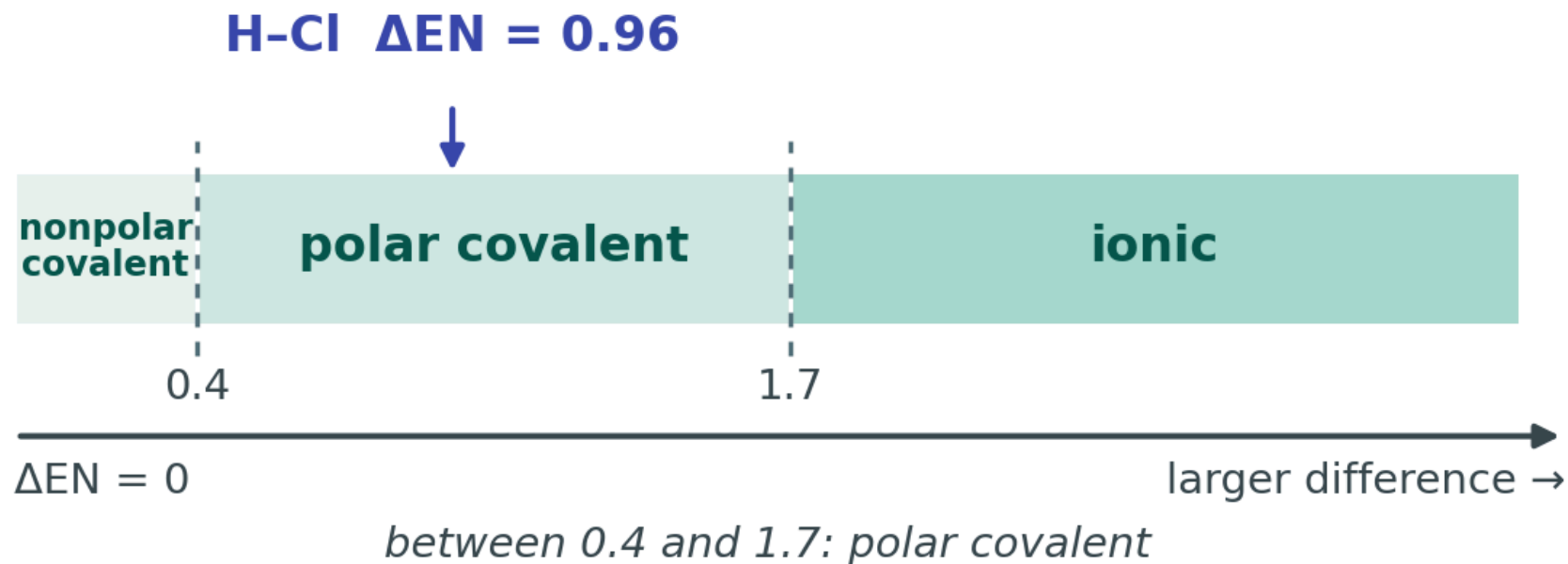
H-Cl → polar covalent, δ^- on Cl
Cl (3.16) pulls harder than H (2.20) · the shared pair shifts toward Cl · H is δ^+

 **Sense check:** The partial negative marks the more electronegative atom, the one that pulls the pair closer. Marking hydrogen reversed the direction the electrons shift.

✓ Worked example 2: the bond on the continuum

$$\text{H-Cl: } \Delta\text{EN} = 3.16 - 2.20 = 0.96$$

electronegativity: H = 2.20 · Cl = 3.16 · found: polar covalent, δ^- on Cl



🧠 **Sense check:** 0.96 lands inside the middle band, between the 0.4 and 1.7 guidelines: unequal sharing.

Take-home: δ^- marks the stronger pull

H-Cl $\rightarrow \delta^-$ on Cl

Cl = 3.16 pulls harder than H = 2.20 · the shared pair shifts toward Cl

subtract backwards: $2.20 - 3.16 = -0.96$

the negative sign points from δ^+ toward δ^- ; it never puts the partial negative on H

The partial negative sits on the more electronegative atom. The size of the difference sets how polar the bond is; the direction of the pull sets which atom is δ^- .

Your turn: the Na-Cl bond

Na-Cl (sodium chloride)

electronegativity: Na = 0.93 · Cl = 3.16

$$\Delta\text{EN} = 3.16 - 0.93 = \underline{\hspace{1cm}} \rightarrow \underline{\hspace{1cm}}$$

Compute the difference, place it on the continuum, and name the bond.

$$\Delta\text{EN} = 3.16 - 0.93 = \mathbf{2.23} \rightarrow \text{a large difference: ionic}$$

Na-Cl → ionic

a metal and a nonmetal · Cl pulls the electron off Na completely · Na⁺ and Cl⁻

Where this goes wrong

Putting δ^- on the wrong atom. Shared electrons shift toward the more electronegative atom. In H-Cl, Cl (3.16) pulls harder than H (2.20), so δ^- sits on Cl. Reading the trend backwards flips the charge onto H; subtracting $2.20 - 3.16 = -0.96$ and dropping the sign does the same.

Calling a real difference nonpolar. Equal sharing happens only when the electronegativities match, as in Cl-Cl. An N-H difference of $3.04 - 2.20 = 0.84$ is not zero, so the sharing is unequal: polar covalent, not nonpolar.

Calling a moderate difference ionic. Two nonmetals with a difference near 1 still share the pair: polar covalent. Full transfer takes a large gap, usually a metal bonded to a nonmetal.

Reading the boundary as a wall. H-F has a difference of 1.78, past the 1.7 guideline, yet HF is a molecular gas that shares its pair: polar covalent. Two nonmetals share; the number guides, the metal-versus-nonmetal test decides at the edge.

Practice 1: the S-O bond

S-O (a sulfur-oxygen bond)

electronegativity: S = 2.58 · O = 3.44 · wanted: bond type and δ -

How is the S-O bond best classified?

- A. Nonpolar covalent, the two atoms share the pair equally
- B. Polar covalent, with S carrying the partial negative charge (δ -)
- C. Polar covalent, with O carrying the partial negative charge (δ -)
- D. Ionic, O pulls an electron completely away from S

✓ Practice 1 answer: C

S-O (a sulfur-oxygen bond)
 electronegativity: S = 2.58 · O = 3.44

$$\Delta\text{EN} = 3.44 - 2.58 = \mathbf{0.86} \rightarrow \text{moderate: polar covalent, } \delta^- \text{ on O} \rightarrow \mathbf{\text{answer C}}$$

B put δ^- on the wrong atom: O (3.44) pulls harder than S (2.58), so the pair shifts toward O; subtracting backwards, $2.58 - 3.44 = -0.86$, only flips the sign, not the direction. A ignored the difference: 0.86 is not zero, so the sharing is unequal, not nonpolar. D read sharing as transfer: two nonmetals 0.86 apart still share the pair, polar covalent, not ionic.

 **Sense check:** A moderate difference means unequal sharing, and the pair sits closer to the more electronegative atom. O is δ^- , S is δ^+ .



Worked example 3: the H-F bond

H-F (hydrogen fluoride)

electronegativity: H = 2.20 · F = 3.98 · wanted: bond type

Hydrogen fluoride is a gas that dissolves in water to make an acid.

A common first answer: the difference clears 1.7, so call the bond ionic. Test it.

✓ Worked example 3: solution

H-F (hydrogen fluoride)
electronegativity: H = 2.20 · F = 3.98

Step 1 · Find the electronegativity difference

$$\Delta\text{EN} = 3.98 - 2.20 = 1.78$$

Step 2 · Place it on the continuum

Step 3 · Name the bond and mark the charges

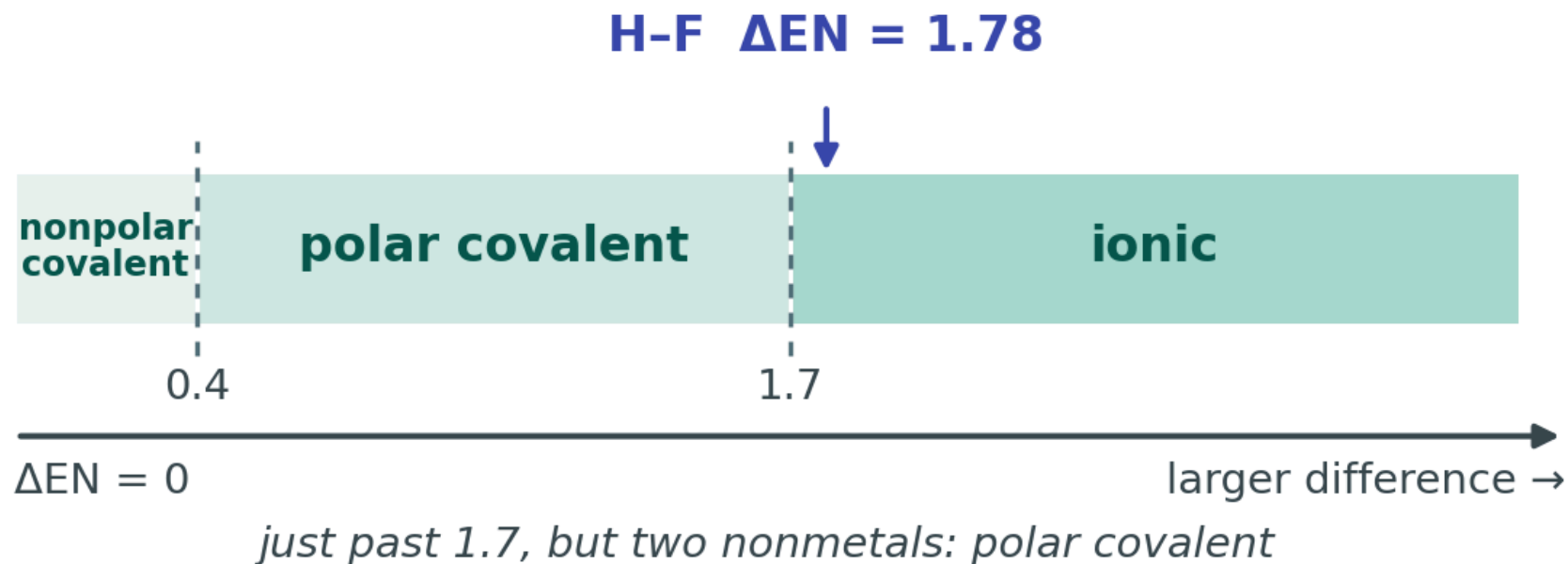
H-F → polar covalent, δ^- on F
two nonmetals · the pair stays shared but pulled far toward F · H is δ^+

 **Sense check:** HF is a molecular gas, not a lattice of ions. The most lopsided sharing among the common bonds is still sharing, not transfer.

✓ Worked example 3: the bond on the continuum

$$\text{H-F: } \Delta\text{EN} = 3.98 - 2.20 = 1.78$$

electronegativity: H = 2.20 · F = 3.98 · found: polar covalent, δ^- on F



 **Sense check:** Just past the 1.7 guideline, yet H and F are both nonmetals: the pair stays shared, polar covalent.

Practice 2: the most polar bond



electronegativity: B = 2.04 · F = 3.98 · Si = 1.90 · Cl = 3.16 · O = 3.44 · wanted: the most polar bond and its δ^- atom

Which of these bonds is the most polar, and which atom in it carries the partial negative charge?

- A. Si-Cl, with Cl carrying δ^-
- B. O-F, with F carrying δ^-
- C. B-F, with B carrying δ^-
- D. B-F, with F carrying δ^-

✓ Practice 2 answer: D



electronegativity: B = 2.04 · F = 3.98 · Si = 1.90 · Cl = 3.16 · O = 3.44

B-F: $3.98 - 2.04 = 1.94$ · Si-Cl: $3.16 - 1.90 = 1.26$ · O-F: $3.98 - 3.44 = 0.54$

→ B-F, δ^- on F → **answer D**

A dropped B-F because 1.94 clears 1.7 and settled for Si-Cl at 1.26; a bigger difference only makes a bond more polar, and with no metal in it B-F still shares its pair. B picked the two most electronegative atoms; polarity comes from the difference, and $3.98 - 3.44 = 0.54$ is the smallest of the three. C put δ^- on the wrong atom: F (3.98) pulls harder than B (2.04); $2.04 - 3.98 = -1.94$ flips only the sign.

 **Sense check:** The biggest difference makes the most polar bond, and 1.7 is a guideline, not a wall: with no metal in the bond, a big difference is still sharing. F is δ^- .



Check yourself

1. A C-O bond has electronegativities C = 2.55 and O = 3.44. Find the difference, classify the bond, and mark the atom that carries δ^- .
2. Two atoms form a bond with a difference of 0. What kind of bond is it, and where do the partial charges sit?

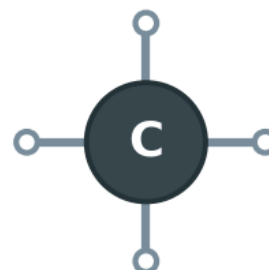
A bond type tells you how one shared pair of electrons is held. A whole molecule is built from several bonds arranged in space, and drawing that arrangement is the next step: Lewis structures.

2 · Lewis Structures

Count a molecule's valence electrons and draw its Lewis structure, placing every electron as a bond or a lone pair.

Parts with a fixed number of connectors

each part has its own number of connectors



one connector two connectors three connectors four connectors



snap two H onto O's two connectors: you build only what the connectors allow

A model kit builds molecules. Each part has connectors: H one, O two, N three, C four. HONC 1-2-3-4. You build only what they allow.

Every valence electron is accounted for

A molecule owns a fixed pool of valence electrons. A Lewis structure places every one of them as bonds or lone pairs. None is invented; none is lost.

H₂O: 8 valence electrons = 4 in bonds + 4 in lone pairs
2 O-H bonds (4) · 2 lone pairs on O (4) · every electron accounted for ✓

Valence electrons come from the group number

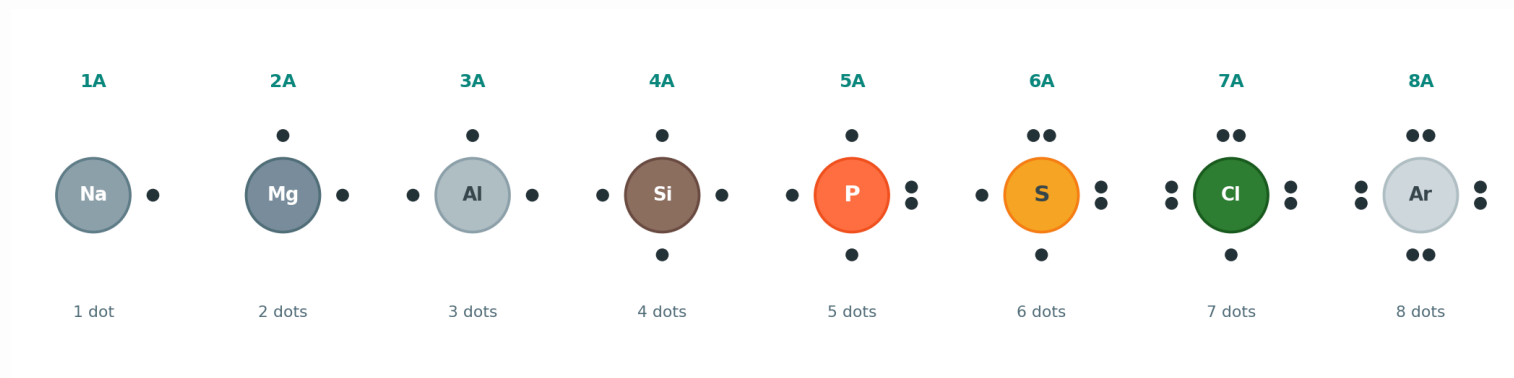
Only the outermost electrons, the valence electrons, form bonds. For a main-group atom their count is the ones digit of the group number.

group 1	group 2	group 13	group 14	group 15	group 16	group 17	group 18
							
1 e ⁻	2 e ⁻	3 e ⁻	4 e ⁻	5 e ⁻	6 e ⁻	7 e ⁻	8 e ⁻

the ones digit of the group number gives the valence-electron count

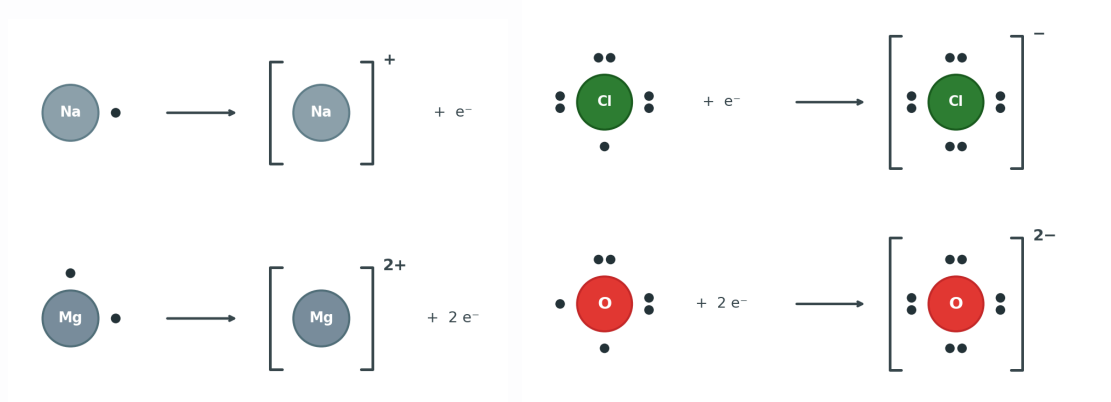
The Lewis symbol: valence electrons as dots

A Lewis symbol writes the element with one dot per valence electron. Dots sit singly on the four sides first, then pair. Eight dots is a full shell.



Ions in Lewis symbols

A metal atom empties its valence shell to form a cation: the dots leave with the electrons. A nonmetal gains electrons until eight dots surround it: an anion. Brackets around the symbol, charge outside.



Your turn: the magnesium dot symbol

Mg: group 2 → 2 valence electrons

12 electrons in the atom · 10 core stay hidden · 2 valence become dots

Fill the blanks, then write the symbol.

step	count
dots around Mg	
paired or single	

·Mg· · 2 dots, single, on two different sides

dots sit singly on the four sides first · pairing starts only at the fifth dot



Practice 1: the dot symbol for phosphorus

P: group 15 → 5 valence electrons

the ones digit of the group number · dots sit singly on the four sides first, then pair

Which dot symbol is correct for a phosphorus atom?

- A. P with 5 dots: one pair and three single dots
- B. P with 3 dots: one on each of three sides
- C. P with 15 dots surrounding the symbol
- D. P with 5 dots: two pairs and one single dot

✓ Practice 1 answer: A

P: group 15 → ones digit 5 → 5 dots → answer A

four dots fill the four sides singly · the fifth makes one pair · one pair + three single dots

B counted the period: P sits in period 3, but dots come from the group, $15 \rightarrow 5$. C drew the whole group number as 15 dots; only the 5 valence electrons appear. D paired too early: dots fill all four sides singly before any pair forms, so 5 dots make one pair, not two.

 **Sense check:** The three single dots mark where phosphorus bonds: three bonds and one lone pair, exactly the NF_3 pattern.



Practice 2: the dot symbol for the sulfide ion

S^{2-} : $6 + 2 = 8$ electrons shown as dots

S is group 16 $\rightarrow$ 6 valence electrons $\cdot$ the $2-$ charge adds 2 more

Which drawing shows the sulfide ion correctly?

- A. [S with 8 dots] $^{2-}$: brackets around the symbol, charge outside
- B. S with 8 dots, no brackets and no charge
- C. [S with 6 dots] $^{2-}$: the six dots sulfur started with
- D. [S with 4 dots] $^{2+}$: two dots removed, a positive charge

✓ Practice 2 answer: A

S^{2-} : $6 + 2 = 8$ dots · brackets, $2-$ outside → answer A
a nonmetal gains electrons until eight dots surround it

B drew the dots but dropped the brackets and charge; eight dots on a plain S claims a neutral atom, and neutral sulfur has 6. C pasted the charge onto the original 6 dots; the $2-$ exists because 2 electrons arrived, so $6 + 2 = 8$ dots must show. D ran the transfer backwards: $6 - 2 = 4$ dots with $2+$ describes a cation, and nonmetals gain.

🧠 **Sense check:** Charge and dot count move together: $2-$ means 2 extra dots, and the brackets say the whole package carries the charge.

The electron pool

Add every atom's valence electrons. For an ion, adjust: add one electron per negative charge, subtract one per positive.

$$\text{CO}_2: 4 + 2(6) = 16$$

1 C (group 14 $\rightarrow$ 4) $\cdot$ 2 O (group 16 $\rightarrow$ 6) $\cdot$ neutral, no adjustment

$$\text{OH}^-: 6 + 1 + 1 = 8 \cdot \text{NH}_4^+: 5 + 4(1) - 1 = 8$$

OH^- adds 1 for the 1 $-$ charge $\cdot$ NH_4^+ subtracts 1 for the 1 $+$ charge

Bonds, lone pairs, and the octet

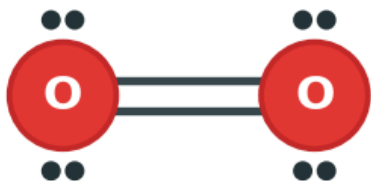
A shared pair drawn as a line is a bond: one pair single, two double, three triple. Each main-group atom aims for eight electrons, an octet. Hydrogen aims for two, a duet.

a shared pair is a bond



one shared pair = single bond

each H reaches a duet (2)



two shared pairs = double bond

each O reaches an octet (8)



three shared pairs = triple bond

each N reaches an octet (8)



The method

1. **Count the valence electrons.**
2. **Draw the skeleton:** least electronegative atom centered, never H; single bonds.
3. **Complete the outer octets:** lone pairs; H a duet.
4. **Place leftovers on the center.**
5. **Form multiple bonds** if the center lacks an octet.



Worked example 1: water

Step 1 • Count the valence electrons

$$\text{H}_2\text{O}: 2(1) + 6 = 8$$

2 H (group 1 $\rightarrow$ 1) $\cdot$ 1 O (group 16 $\rightarrow$ 6) $\cdot$ 8 valence electrons to place

Oxygen is less electronegative than hydrogen, and H is never central. Draw the Lewis structure.

✓ Worked example 1: solution



8 valence electrons to place

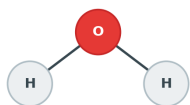
Step 2 • Draw the skeleton



2 single bonds use $2 \times 2 = 4$ electrons • 4 of 8 placed

Step 3 • Complete the outer octets

The outer atoms are hydrogen; a single bond already fills each H's duet.



🧠 **Sense check:** Four electrons fill the two O-H bonds and each hydrogen has its duet. Four remain for oxygen.

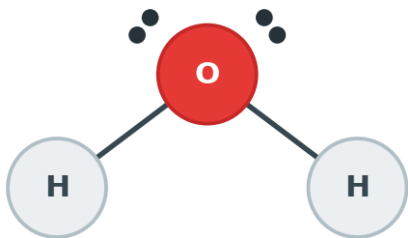
✓ Worked example 1: the oxygen lone pairs

Four electrons sit in the two O-H bonds; the outer hydrogens are done. Four remain.

Step 4 • Place leftovers on the center

$8 - 4 = 4$ electrons left $\rightarrow$ 2 lone pairs on O

O: 2 bonds (4) + 2 lone pairs (4) = 8 $\rightarrow$ octet ✓



🧠 **Sense check:** All 8 valence electrons placed: 4 in bonds, 4 in lone pairs. O has an octet; each H a duet.

✓ Worked example 1: the electron budget

H_2O : 8 valence electrons = 4 in bonds + 4 in lone pairs on O
steps 1 to 4 used · step 5 not needed

H_2O : 8 valence electrons

1 · Count the valence electrons		8 to place
2 · Draw the skeleton		2 O-H bonds: 4 placed · 4 left
3 · Complete the outer octets		each H has its duet: 0 added · 4 left
4 · Place leftovers on the center		2 lone pairs on O · 0 left
5 · Form multiple bonds		not needed: O has an octet

in bonds
 lone pairs on outer atoms
 lone pairs on the center
 lone pairs moved into bonds

 **Sense check:** The pool is used up at step 4 and oxygen has its octet, so step 5 never runs. ✓



Worked example 2: nitrogen trifluoride

Step 1 • Count the valence electrons



given: 1 nitrogen, 3 fluorine • wanted: the electron pool, then the structure

A common first attempt counts one group number per element. Find the pool, then draw the structure.

✓ Worked example 2: solution

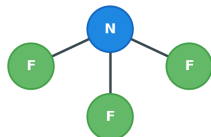
Step 1 • Count the valence electrons

NF_3 : $5 + 3(7) = 26$ • one-per-element $5 + 7 = 12 \times$
each F brings its own 7: multiply by the subscript • 26 to place

Step 2 • Draw the skeleton

Nitrogen is least electronegative, so it takes the center, with a single bond to each F.

3 N-F single bonds: $3 \times 2 = 6$ electrons
6 of 26 placed



 **Sense check:** Six of the 26 electrons fill the three N-F bonds. Twenty remain for the fluorine octets and the center.

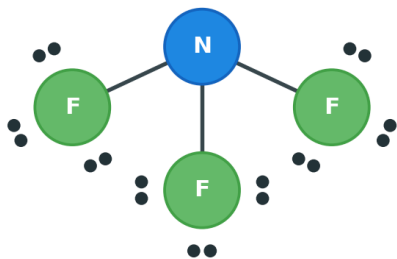
✓ Worked example 2: the fluorine octets

Six electrons sit in the three N-F bonds. Each fluorine still needs three lone pairs.

Step 3 • Complete the outer octets

each F: 3 lone pairs $\rightarrow 3 \times 6 = 18$ electrons

$6 + 18 = 24$ of 26 placed • every F has its octet



🧠 **Sense check:** Twenty-four of the 26 electrons are placed and every fluorine has its octet. Two remain for the center.

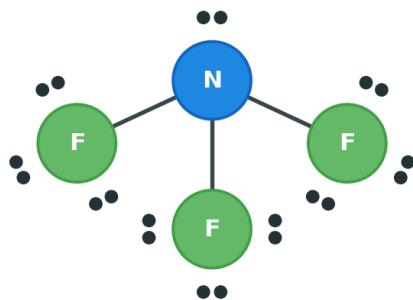
✓ Worked example 2: the nitrogen lone pair

Six electrons sit in the three N-F bonds and eighteen fill the fluorine octets. Two remain.

Step 4 • Place leftovers on the center

$$26 - 6 - 18 = 2 \rightarrow 1 \text{ lone pair on N}$$

$$\text{N: } 3 \text{ bonds (6)} + 1 \text{ lone pair (2)} = 8 \rightarrow \text{octet } \checkmark$$



🧠 **Sense check:** All 26 valence electrons placed; every F and the N reaches an octet. One per element stops at 12 and leaves the structure short.

✓ Worked example 2: the electron budget

NF₃: 26 valence electrons = 6 in bonds + 18 on the F atoms + 2 on N
 steps 1 to 4 used · step 5 not needed

NF₃: 26 valence electrons

- 1 · Count the valence electrons  26 to place
 - 2 · Draw the skeleton  3 N-F bonds: 6 placed · 20 left
 - 3 · Complete the outer octets  3 lone pairs on each F: 18 · 2 left
 - 4 · Place leftovers on the center  1 lone pair on N · 0 left
 - 5 · Form multiple bonds  *not needed: N has an octet*
-  in bonds
  lone pairs on outer atoms
  lone pairs on the center
  lone pairs moved into bonds

 **Sense check:** The same four steps as water. The outer octets take the largest share: 18 of 26. ✓

Take-home: count every atom, not every element

$$\text{NF}_3: 5 + 3(7) = 26 \checkmark \quad \cdot \quad 5 + 7 = 12 \times$$

three fluorine atoms bring $3 \times 7 = 21$ electrons, not 7

Each atom brings its own valence electrons. Multiply each element's group number by its subscript in the formula, then add.

Your turn: methane, CH₄

$$\text{CH}_4: 4 + 4(1) = 8$$

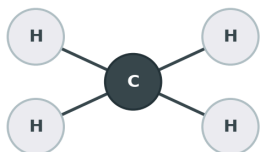
1 C (group 14 $\rightarrow$ 4) $\cdot$ 4 H (group 1 $\rightarrow$ 1) $\cdot$ 8 valence electrons

Carbon is the central atom. Fill the blanks.

step	count
electrons in the 4 C-H bonds	<u> </u>
electrons left for lone pairs	<u> </u>

$$4 \text{ C-H bonds: } 4 \times 2 = 8 \cdot \text{ leftover } 8 - 8 = 0$$

C: 4 bonds (8) = octet $\checkmark$ $\cdot$ every H a duet $\cdot$ no lone pairs



⚠ Where this goes wrong

Counting each element once. For NF_3 , adding one group number per element gives $5 + 7 = 12$. Every atom brings its own: $5 + 3(7) = 26$. Multiply each group number by its subscript.

Counting all the electrons. CO_2 holds $6 + 2(8) = 22$ electrons in total, but only the valence electrons are drawn: $4 + 2(6) = 16$. Core electrons stay out of the structure.

Ignoring the ion's charge. NH_4^+ carries a 1+ charge, so subtract one electron: $5 + 4(1) - 1 = 8$, not 9. A negative ion adds electrons; a positive ion removes them.

Leaving the center short. With only single bonds, the carbon in CO_2 has 4 electrons, not 8. When the center lacks an octet, form double or triple bonds.

Practice 3: oxygen difluoride



fluorine is the most electronegative atom, so oxygen is the central atom

How many valence electrons must appear in the Lewis structure of OF_2 ?

- A. 18
- B. 13
- C. 26
- D. 20

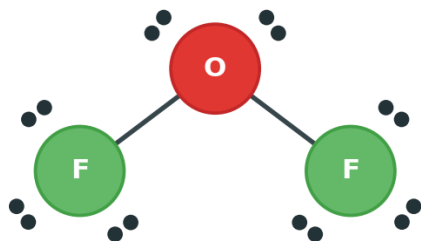
✓ Practice 3 answer: D

$$\text{OF}_2: 6 + 2(7) = 20 \rightarrow \text{answer D}$$

1 O (group 16 $\rightarrow$ 6) $\cdot$ 2 F (group 17 $\rightarrow$ 7) $\cdot$ neutral, no adjustment

B counted each element once: $6 + 7 = 13$. C counted every electron, core included: $8 + 2(9) = 26$. A dropped a pair: $20 - 2 = 18$.

🧠 **Sense check:** 20 valence electrons place as 2 O-F bonds (4) and 8 lone pairs (16): $4 + 16 = 20$, every atom an octet.





Worked example 3: carbon dioxide

Step 1 • Count the valence electrons

$$\text{CO}_2: 4 + 2(6) = 16$$

1 C (group 14 $\rightarrow$ 4) $\cdot$ 2 O (group 16 $\rightarrow$ 6) $\cdot$ 16 valence electrons to place

Carbon is least electronegative, so it takes the center. Draw the Lewis structure.

✓ Worked example 3: solution

Step 2 • Draw the skeleton

Carbon centered, a single bond to each oxygen.

O-C-O: 2 single bonds use $2 \times 2 = 4$ electrons
4 of 16 placed

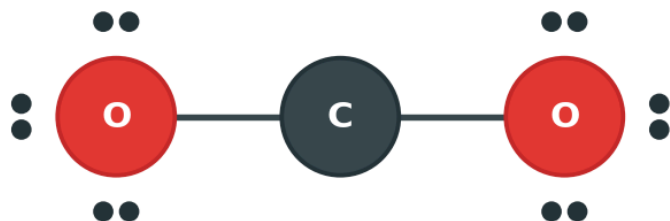


🧠 **Sense check:** Four of the 16 electrons fill the two C-O bonds. Twelve remain for the outer octets.

✓ Worked example 3: the outer octets

Step 3 • Complete the outer octets

each O gets 3 lone pairs: $2 \times 6 = 12$ electrons
 $4 + 12 = 16$: all placed, but carbon has only 4 (no octet)



Step 4 • Place leftovers on the center

None remain: $16 - 4 - 12 = 0$. The carbon is still two pairs short of an octet.

🧠 **Sense check:** Every electron is placed, yet carbon holds only 4. Single bonds cannot finish this structure.

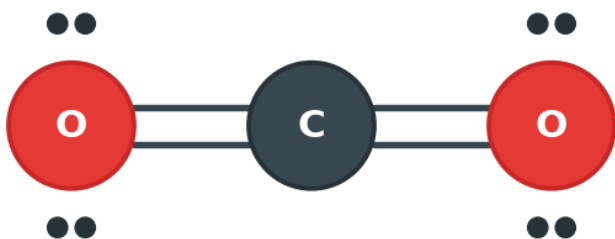
✓ Worked example 3: two double bonds

Carbon sits two pairs short of an octet. Pull one lone pair from each oxygen into a bond.

Step 5 • Form multiple bonds



C: 2 double bonds (8) = octet ✓ · each O: 1 double bond (4) + 2 lone pairs (4) = 8 ✓



🧠 **Sense check:** All 16 valence electrons placed. Single bonds could not give carbon an octet; two double bonds can.

✓ Worked example 3: the electron budget

CO₂: 16 valence electrons = 8 in bonds + 8 in lone pairs on O
all five steps used · step 4 adds nothing: none left

CO₂: 16 valence electrons

1 · Count the valence electrons



2 · Draw the skeleton



3 · Complete the outer octets



4 · Place leftovers on the center



5 · Form multiple bonds



■ in bonds
 ■ lone pairs on outer atoms
 ■ lone pairs on the center
 ■ lone pairs moved into bonds

 **Sense check:** The pool ran out with carbon at 4. Step 5 moves 2 lone pairs into bonds: 4 + 4 = 8 in bonds. No electron is added. ✓

Practice 4: hydrogen cyanide



In the Lewis structure of HCN, how many electrons sit in lone pairs?

- A. 10
- B. 2
- C. 6
- D. 8
- E. 4

✓ Practice 4 answer: B

HCN: $1 + 4 + 5 = 10$ · H-C≡N: 1 single bond (2) + 1 triple bond (6) = 8 shared · $10 - 8 = 2$ in lone pairs → answer B

C: single (2) + triple (6) = 8 ✓ · N: triple (6) + 1 lone pair (2) = 8 ✓ · H: duet ✓

A reported the whole pool of 10, not the lone-pair share. D counted the shared electrons, $2 + 6 = 8$, instead of the ones left over. C stopped at single bonds: $10 - 4 = 6$ in lone pairs, which leaves carbon with only 4 electrons. E stopped at a double bond: $2 + 4 = 6$ shared, $10 - 6 = 4$ in lone pairs, which leaves carbon with only 6.

 **Sense check:** Two of nitrogen's pairs move into the bond so carbon reaches eight. One lone pair, 2 electrons, is all that stays outside a bond.

Check yourself

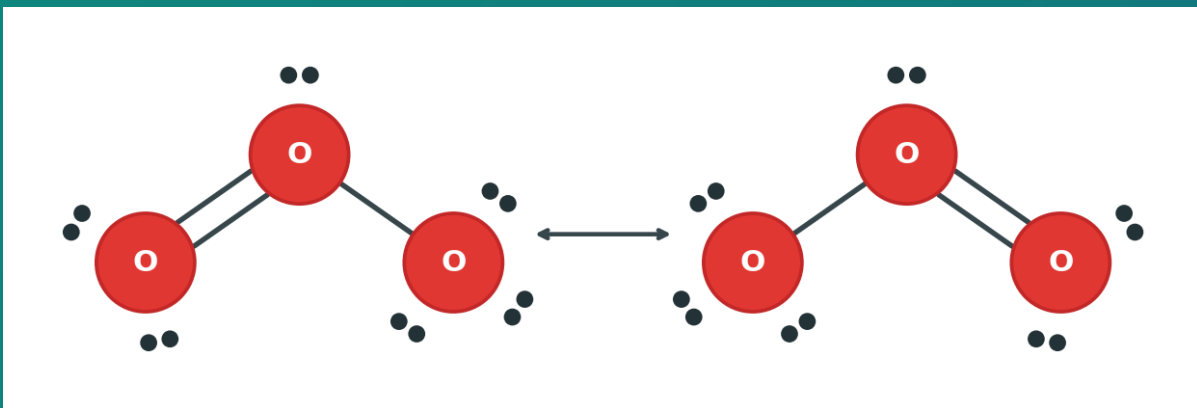
1. Count the valence electrons in NH_3 , and name the central atom.
2. In CO_2 , why do single bonds fail and double bonds succeed?

The octet rule builds most structures, not all: boron can settle at six electrons, and heavier central atoms can hold more than eight. Before those exceptions, every finished structure must survive a bookkeeping check: each atom's share of the electrons, counted against the pool.

3 · Resonance & Formal Charge

Draw every resonance structure of a molecule or ion, assign formal charges, and use them to pick the best Lewis structure.

One molecule, two correct drawings



High in the stratosphere, ozone absorbs ultraviolet light. Its Lewis structure can be drawn two equally correct ways. Measured bond lengths match neither drawing: both bonds are identical.

Delocalized electrons defeat a single drawing

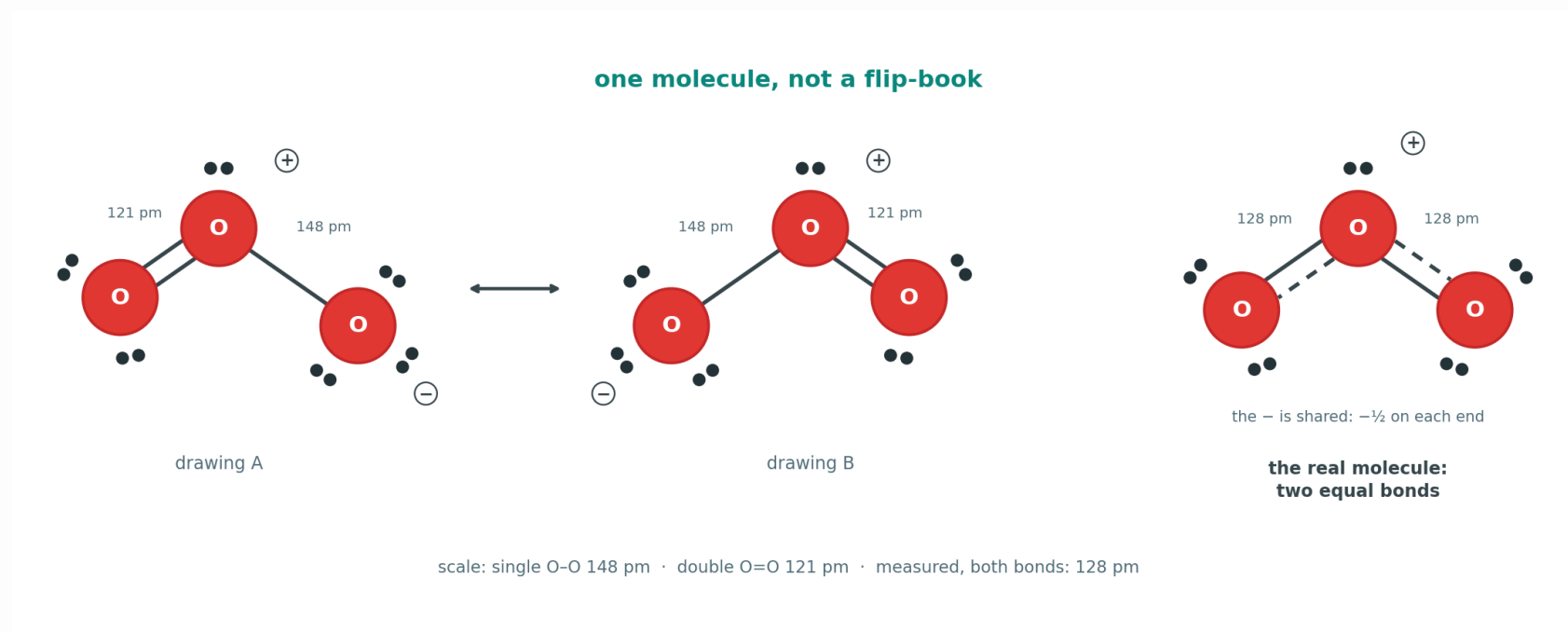
A Lewis structure pins every electron pair to one place. In ozone the fourth bonding pair belongs to both O-O links at once. No single drawing shows that, so two share the job.

O_3 : $3(6) = 18$ valence electrons · one double bond, drawable on either side

$\text{O}=\text{O}-\text{O} \leftrightarrow \text{O}-\text{O}=\text{O}$ · atoms fixed · only electrons move

Resonance structures and the blend

Resonance structures differ only in where multiple bonds and lone pairs sit. The real molecule is one blend of all of them: every bond equal, between single and double.



O_3 , NO_2^- , CO_3^{2-} , NO_3^- : equivalent resonance structures
connect the drawings with a double-headed arrow $\leftrightarrow$ · never move an atom

Charges you already count

$$\text{Na}^+: 1 - 0 = +1 \quad \cdot \quad \text{Cl}^-: 7 - 8 = -1$$

charge = valence electrons brought – electrons held · Na: 1 and 0 · Cl: 7 and 8

Monatomic ions already follow this count. Sodium gives up its one valence electron: +1. Chlorine gains one to hold eight: –1. Formal charge runs the same count on every atom inside a Lewis structure.

Formal charge tracks electron ownership

Formal charge compares the electrons an atom holds in the structure with the electrons the free atom brings. Each atom owns its nonbonding electrons and half of every bond it makes.

$$\text{FC} = \text{valence} - \text{nonbonding} - \frac{1}{2}(\text{bonding})$$

valence: the free atom's count · nonbonding: lone-pair electrons · bonding: electrons in its bonds

memory hook: valence minus dots minus lines

each lone-pair electron is one dot · each bond line counts 1, half of its 2 electrons

The formal charges sum to the overall charge

Add every atom's formal charge. A neutral molecule sums to zero; an ion sums to its charge. A different total means a miscount somewhere in the structure.

$$\text{O}_3: (-1) + (+1) + 0 = 0 \checkmark \quad \cdot \quad \text{CO}_3^{2-}: 0 + 0 + (-1) + (-1) = -2 \checkmark$$

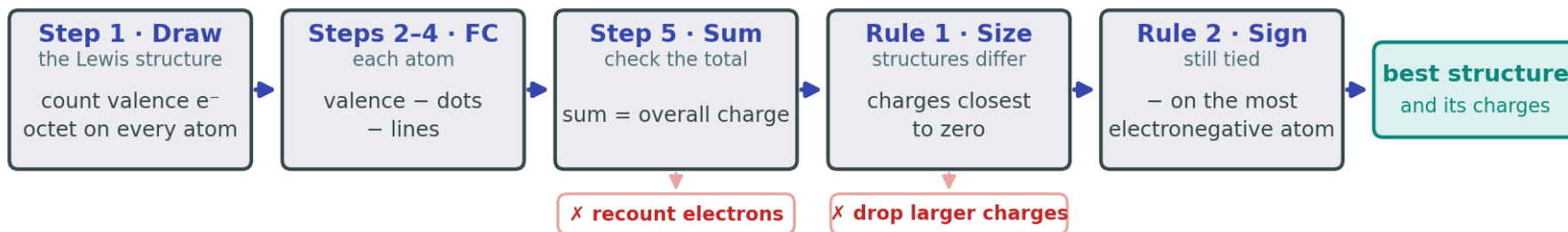
sum equals the overall charge, or the structure has an error



The method

1. Draw the full Lewis structure.
2. Count the atom's nonbonding electrons.
3. Count half the atom's bonding electrons.
4. Subtract both from the valence electrons.
5. Check the sum against the overall charge.

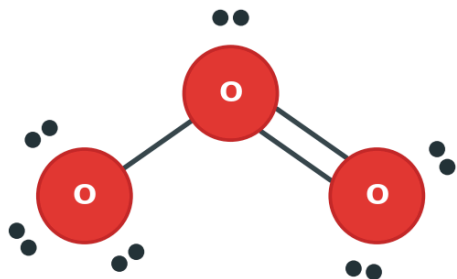
Count every atom, check the sum, then choose. Rules 1 and 2 only when structures differ.



Worked example 1: formal charges in ozone

Step 1 · Draw the full Lewis structure

O_3 : $3(6) = 18$ valence electrons, all placed
given: the structure below · wanted: the formal charge on each oxygen



Assign a formal charge to each oxygen.

✓ Worked example 1: solution

$$\text{O}-\text{O}=\text{O} \cdot \text{FC} = \text{valence} - \text{nonbonding} - \frac{1}{2}(\text{bonding})$$

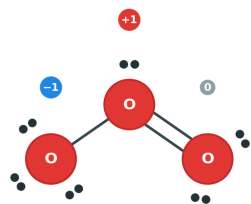
Step 2 · Count the atom's nonbonding electrons

Step 3 · Count half the atom's bonding electrons

Step 4 · Subtract both from the valence electrons

$$6 - 6 - 1 = -1 \quad \cdot \quad 6 - 2 - 3 = +1 \quad \cdot \quad 6 - 4 - 2 = 0$$

single-bonded O: -1 · central O: $+1$ · double-bonded O: 0



 **Sense check:** The negative charge sits on the oxygen holding the most lone pairs. Bonds count half; lone pairs count in full.

✓ Worked example 1: the sum check

O-O=O: formal charges $-1, +1, 0$
 single-bonded O · central O · double-bonded O

Step 5 · Check the sum against the overall charge

$(-1) + (+1) + 0 = 0$
 ozone is neutral · the formal charges sum to 0 ✓

 **Sense check:** Any total other than zero would mean a dropped electron or a miscounted bond somewhere in the structure.



Worked example 2: the carbonate ion

Step 1 · Draw the full Lewis structure

$$\text{CO}_3^{2-}: 4 + 3(6) + 2 = 24 \text{ valence electrons}$$

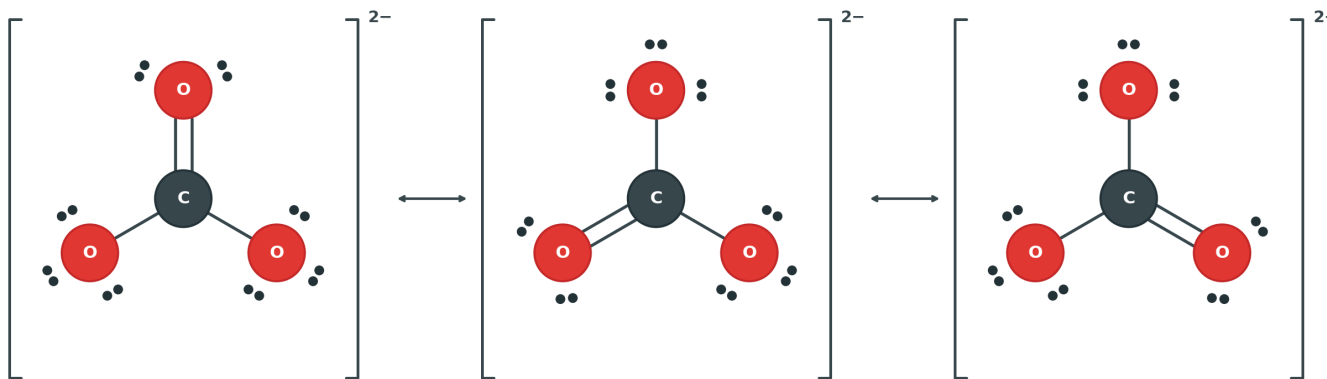
the 2– charge adds 2 electrons · one C=O double bond completes carbon's octet

The double bond can sit on any of the three oxygens. Draw every resonance structure, then assign the formal charges.

✓ Worked example 2: the three structures

CO_3^{2-} : 24 valence electrons · one double bond, three equivalent positions
atoms never move · only the double bond and lone pairs shift

The double bond takes each position in turn.



 **Sense check:** Three drawings, one ion. The real carbonate has three equal C-O bonds, each between a single and a double.

✓ Worked example 2: formal charges

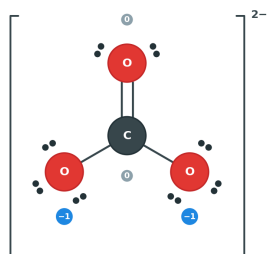
one structure of CO_3^{2-} : C with 4 bonds · one double-bonded O · two single-bonded O

Step 2 · Count the atom's nonbonding electrons

Step 3 · Count half the atom's bonding electrons

Step 4 · Subtract both from the valence electrons

C: $4 - 0 - 4 = 0$ · O(=): $6 - 4 - 2 = 0$ · O(-): $6 - 6 - 1 = -1$
the two single-bonded oxygens each carry -1



 **Sense check:** Bonds count half, lone pairs in full. The negative charges land on the single-bonded oxygens.

✓ Worked example 2: the sum check

one structure of CO_3^{2-} : C 0 · O(=) 0 · two O(-) at -1 each
formal charges from the previous step

Step 5 · Check the sum against the overall charge

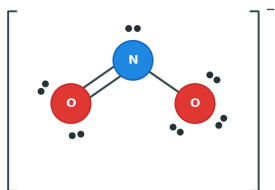
$$0 + 0 + (-1) + (-1) = -2$$

matches the ion's 2- charge ✓

 **Sense check:** Every resonance structure gives the same set of formal charges; only their positions rotate with the double bond.

Your turn: the nitrite ion

NO_2^- : $5 + 2(6) + 1 = 18$ valence electrons
 structure: $\text{O}=\text{N}-\text{O}$ in brackets · N keeps one lone pair



atom	formal charge
N: $5 - 2 - 3$	_____
single-bonded O: $6 - 6 - 1$	_____

$\text{N: } 5 - 2 - 3 = 0 \cdot \text{O}(-): 6 - 6 - 1 = -1 \cdot \text{sum: } 0 + 0 + (-1) = -1 \checkmark$
 double-bonded O: 0 · the sum matches the 1– charge

⚠ Where this goes wrong

Treating resonance as motion. The molecule does not flip between the drawings. There is one real structure: the blend, with equal intermediate bonds.

Counting all the bonding electrons. For ozone's central O: $6 - 2 - 6 = -2$ ✗. An atom owns half of its bonds: $6 - 2 - 3 = +1$ ✓.

Moving atoms between structures. Rearranged atoms make a different compound. Resonance structures move electrons only.

Skipping the sum check. The formal charges must add to the overall charge; any other total means a dropped electron or a miscounted bond.

Practice 1: the ammonium ion

NH_4^+ : N bonded to 4 H, no lone pair on N

$5 + 4(1) - 1 = 8$ valence electrons · all four pairs sit in N-H bonds

What is the formal charge on the nitrogen atom?

- A. -3
- B. -1
- C. +1
- D. 0

✓ Practice 1 answer: C

$$\text{N: } 5 - 0 - 4 = +1 \rightarrow \text{answer C}$$

valence 5 · nonbonding 0 · four bonding pairs hold 8, half is 4 · each H: $1 - 0 - 1 = 0$ · sum +1 ✓

 **Sense check:** A nitrogen with four bonds and no lone pair owns one electron fewer than its five: +1, the charge of the whole ion.

the ammonium ion, NH_4^+



Choosing between non-equivalent structures

When two valid structures are not equivalent, formal charge decides. Prefer the structure with charges closest to zero, and put any negative formal charge on the most electronegative atom.

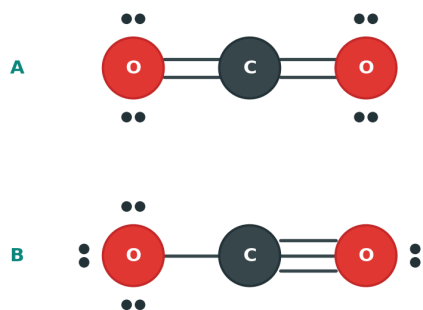
best structure: smallest formal charges · negative FC on the most electronegative atom

Rule 1 · size, then Rule 2 · sign, applied in that order

Worked example 3: two structures for carbon dioxide

CO_2 : $4 + 2(6) = 16$ valence electrons in each structure

A: $\text{O}=\text{C}=\text{O}$ · B: $\text{O}-\text{C}\equiv\text{O}$ · every atom has an octet in both



Both structures place all 16 electrons. Which one better represents CO_2 ?

✓ Worked example 3: solution

$$\text{A: O } 6 - 4 - 2 = 0 \cdot \text{C } 4 - 0 - 4 = 0$$

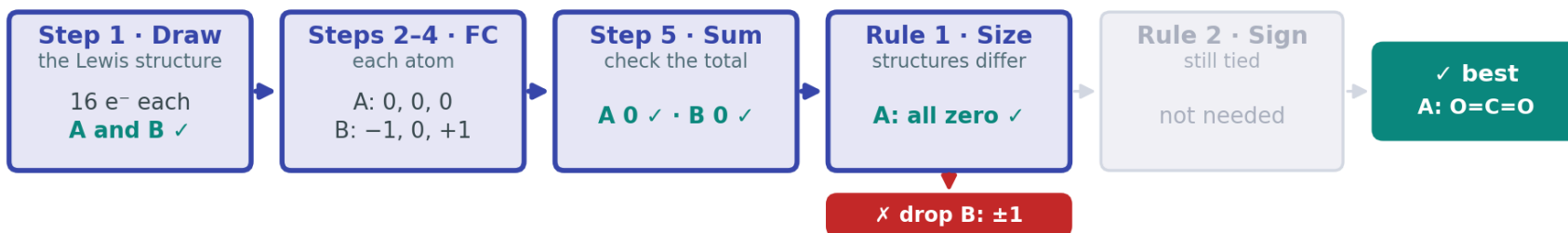
$$\text{B: O(-) } 6 - 6 - 1 = -1 \cdot \text{C } 0 \cdot \text{O(}\equiv\text{) } 6 - 2 - 3 = +1$$

both sums 0 ✓ · B's +1 lands on oxygen, the most electronegative atom

Rule 1 · Size: charges closest to zero

🧠 **Sense check:** Both sums are zero, so size decides: A has no formal charge, B has ± 1 . A wins; CO_2 is $\text{O}=\text{C}=\text{O}$.

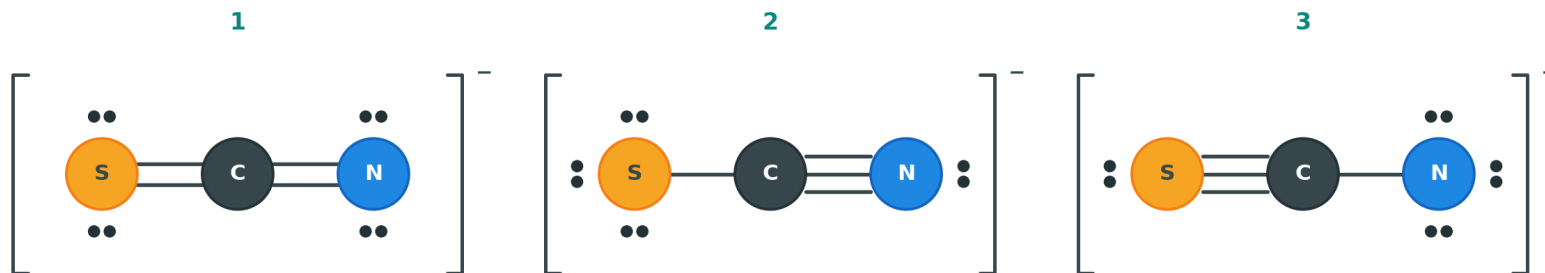
carbon dioxide: A $\text{O}=\text{C}=\text{O}$ or B $\text{O}-\text{C}\equiv\text{O}$



Guided example: the thiocyanate ion

Step 1 · Draw the full Lewis structure

SCN^- : $6 + 4 + 5 + 1 = 16$ valence electrons · skeleton S-C-N
three octet structures, not equivalent · electronegativity: N 3.04 · S 2.58



Assign every formal charge in all three structures. Then choose the best structure.

✓ Guided example: formal charges

SCN⁻: 16 valence electrons · 1: S=C=N · 2: S-C≡N · 3: S≡C-N

wanted: the formal charge on every atom of all three

Step 2 · Count the atom's nonbonding electrons

Step 3 · Count half the atom's bonding electrons

Step 4 · Subtract both from the valence electrons

Step 5 · Check the sum against the overall charge

structure	S	C	N	sum
1: S=C=N	$6 - 4 - 2 = 0$	0	$5 - 4 - 2 = -1$	$0 + 0 - 1 = -1 \checkmark$
2: S-C≡N	$6 - 6 - 1 = -1$	0	$5 - 2 - 3 = 0$	$-1 + 0 + 0 = -1 \checkmark$
3: S≡C-N	$6 - 2 - 3 = +1$	0	$5 - 6 - 1 = -2$	$+1 + 0 - 2 = -1 \checkmark$

 **Sense check:** Carbon stays at 0; the charges shift between S and N as the multiple bonds move. All three match the 1- charge.

✓ Guided example: the best structure

1: $\text{S}=\text{C}=\text{N}$ (0, 0, -1) · 2: $\text{S}-\text{C}\equiv\text{N}$ (-1, 0, 0) · 3: $\text{S}\equiv\text{C}-\text{N}$ (+1, 0, -2)
 SCN^- · formal charges S, C, N · all three sum to -1 ✓ · electronegativity: N 3.04 · S 2.58

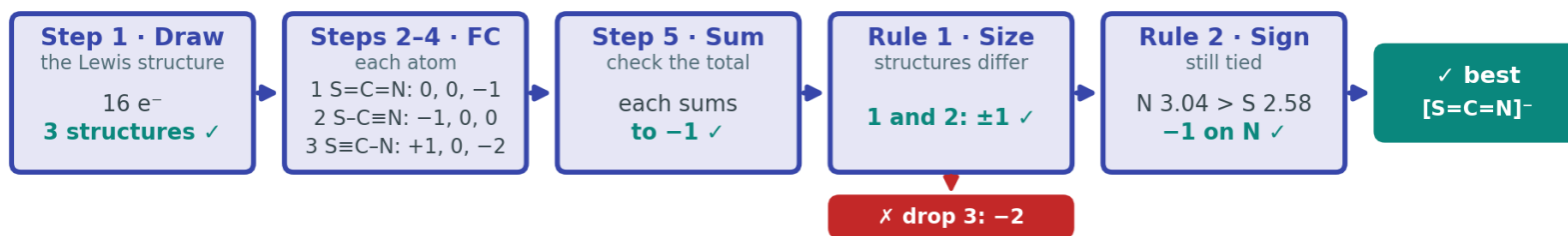
Rule 1 · Size: charges closest to zero

Rule 2 · Sign: negative charge on the most electronegative atom

best: $[\text{S}=\text{C}=\text{N}]^-$ · S 0 · C 0 · N -1
 N 3.04 – S 2.58 = 0.46 · the -1 sits on N

 **Sense check:** Structures 1 and 2 tie on size, so sign decides: N is more electronegative than S, and the -1 belongs on N.

the thiocyanate ion, SCN^-



Practice 2: two structures for N₂O

N₂O: 2(5) + 6 = 16 valence electrons · structure 1: N≡N-O · structure 2: N=N=O
every atom has an octet in both · wanted: the better structure

Which structure better represents N₂O?

- A. Structure 2: every formal charge is zero, as in O=C=O
- B. Structure 1: both carry a -1 and a +1, and this one puts the -1 on oxygen, the more electronegative atom
- C. Structure 2: its -1 sits on the end nitrogen, which holds negative charge better than oxygen does
- D. Both equally: the two are resonance structures of the same molecule, so they count the same

✓ Practice 2 answer: B

structure 1, $\text{N}\equiv\text{N}-\text{O}$: $\text{N } 5 - 2 - 3 = 0 \cdot \text{N } 5 - 0 - 4 = +1 \cdot \text{O } 6 - 6 - 1 = -1$
 sum $0 + 1 - 1 = 0 \checkmark$ · the -1 on oxygen

structure 2, $\text{N}=\text{N}=\text{O}$: $\text{N } 5 - 4 - 2 = -1 \cdot \text{N } 5 - 0 - 4 = +1 \cdot \text{O } 6 - 4 - 2 = 0$
 sum $-1 + 1 + 0 = 0 \checkmark$ · the -1 on nitrogen $\rightarrow$ answer B

A miscounted: the central N holds four bonds and no lone pair in both drawings, $5 - 0 - 4 = +1$, so no structure of N_2O is charge-free. C ran the tiebreaker backwards; a negative formal charge belongs on the most electronegative atom, and oxygen (3.44) beats nitrogen (3.04). D treated non-equivalent drawings as equal partners; their charges differ, so formal charge picks the better one.

 **Sense check:** Same charge sizes in both, so the second rule decides: the -1 goes to oxygen. $\text{N}\equiv\text{N}-\text{O}$ is the better structure.

Practice 3: the cyanate ion



skeleton O-C-N · wanted: the formal charges in the best structure

In the best Lewis structure of OCN^- , what is the formal charge on each atom?

- A. O 0, C 0, N -1
- B. O +1, C 0, N -2
- C. O -2, C +4, N -3
- D. O 0, C 0, N 0
- E. O -1, C 0, N 0

✓ Practice 3 answer: E

$[\text{O}-\text{C}\equiv\text{N}]^-$: O $6 - 6 - 1 = -1$ · C $4 - 0 - 4 = 0$ · N $5 - 2 - 3 = 0 \rightarrow$ answer E
 $4 + 6 + 5 + 1 = 16$ electrons · sum -1 ✓ · ties $[\text{O}=\text{C}=\text{N}]^-$ on size · O $3.44 >$ N 3.04 takes the -1

A is $[\text{O}=\text{C}=\text{N}]^-$: O $6 - 4 - 2 = 0$, N $5 - 4 - 2 = -1$. It ties E on size but puts the -1 on N, as in $[\text{S}=\text{C}=\text{N}]^-$; oxygen outranks nitrogen. B is $[\text{O}\equiv\text{C}-\text{N}]^-$: O $6 - 2 - 3 = +1$, N $5 - 6 - 1 = -2$, which fails Rule 1. C gave oxidation numbers: O -2 , N -3 , so C is $-1 + 2 + 3 = +4$. D sums to 0; an ion's formal charges must sum to -1 .

 **Sense check:** Same skeleton as thiocyanate, opposite answer. Rule 2 asks which end atom is more electronegative: N beats S, but O beats N.



Check yourself

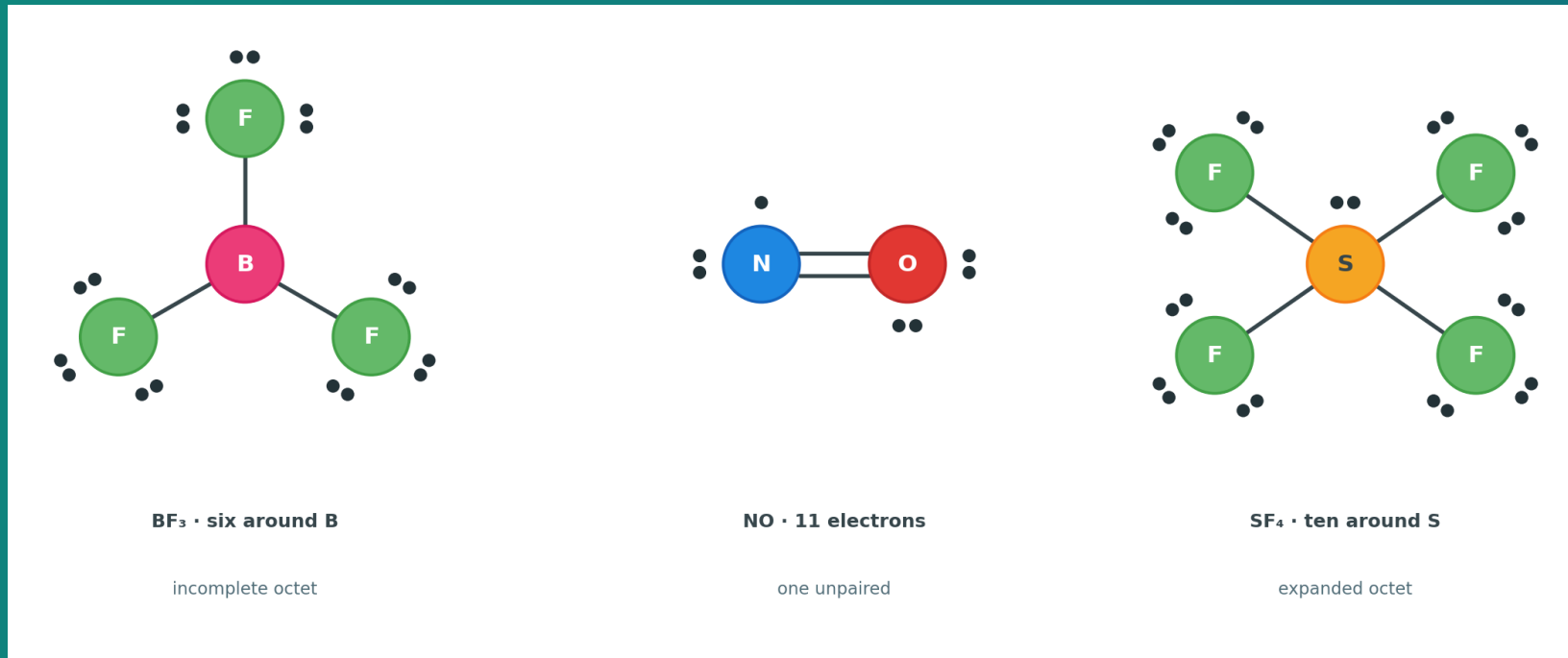
1. Why does a single Lewis structure fail for the carbonate ion?
2. Compute the formal charge on the double-bonded oxygen in ozone.

Formal charge also referees the octet rule itself: some structures are best with fewer than eight electrons on an atom, and heavier central atoms can hold more. Those exceptions are the next topic.

4 • Exceptions to the Octet Rule

Recognize and draw the three exception families: incomplete octets, odd-electron molecules, and expanded octets on period-3-or-lower central atoms.

Three ways real molecules break the octet rule



The octet rule builds most structures, not all. Boron trifluoride stops at six electrons. Nitrogen monoxide carries an odd electron. Sulfur tetrafluoride holds ten.

The octet is a pattern, not a law

Eight electrons fill an atom's s and p valence orbitals. Most main-group atoms reach that count. Three families do not, each for its own reason.

BF₃: six around B · NO: 11 electrons, one unpaired · SF₄: ten around S
incomplete octet · odd-electron molecule · expanded octet

period 3 or lower can expand · C, N, O, F stop at eight
a period-2 atom has four valence orbitals: room for eight, never more

Incomplete octets: beryllium and boron

Boron and beryllium routinely stop short of eight. In BF_3 , boron holds three bonds and six electrons, and the structure is stable as drawn.

BF_3 : $3 + 3(7) = 24$ · B: 3 bonds = 6 electrons
every F keeps its octet · boron does not reach eight

Odd-electron molecules

An odd valence total can never pair into full octets. NO holds 11 electrons; one stays unpaired on nitrogen. The structure shows it as a single dot.

NO: $5 + 6 = 11$ valence electrons

10 pair up · 1 electron remains single · no octet for N

Expanded octets: period 3 and below

A period-2 atom has four valence orbitals, one s and three p: eight electrons at most. Atoms from period 3 down are larger, with more orbitals, so more pairs fit.

PF_5 : 10 around P · SF_4 : 10 around S · ICl_4^- : 12 around I
only period-3-or-lower centers expand · C, N, O, F stop at eight



The method

1. **Count the valence electrons.**
2. **Build the skeleton, then the outer octets.**
3. **Leftovers go on the central atom**, past eight when it sits in period 3 or lower.
4. **Never expand a second-period atom.**



Worked example 1: phosphorus pentafluoride

Step 1 • Count the valence electrons

$$\text{PF}_5: 5 + 5(7) = 40$$

1 P (group 15 $\rightarrow$ 5) $\cdot$ 5 F (group 17 $\rightarrow$ 7) $\cdot$ 40 to place

Five fluorines must bond to one phosphorus. Draw the Lewis structure.

✓ Worked example 1: solution

PF_5 : $5 + 5(7) = 40$ valence electrons · skeleton: five P-F single bonds

Step 2 · Build the skeleton, then the outer octets

5 bonds: $5 \times 2 = 10$ · 5 F $\times$ 3 lone pairs: $5 \times 6 = 30$

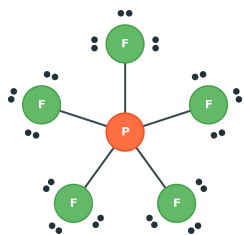
$10 + 30 = 40$ · every electron placed · every F has its octet

🧠 **Sense check:** Ten electrons sit in the five P-F bonds and thirty complete the fluorines. All 40 are placed.

✓ Worked example 1: the expanded octet

Step 3 · Leftovers go on the central atom

$40 - 10 - 30 = 0$ leftover · P: 5 bonds = 10 electrons
phosphorus exceeds eight through its bonds alone · period 3 → allowed



🧠 **Sense check:** Ten electrons around phosphorus: an expanded octet carried by the bonds alone.



Worked example 2: the ICl_4^- ion

Step 1 • Count the valence electrons

$$\text{ICl}_4^-: 7 + 4(7) + 1 = 36$$

the 1– charge adds one electron · 36 to place

Iodine takes the center. Draw the Lewis structure.

✓ Worked example 2: bonds and outer octets

ICl_4^- : $7 + 4(7) + 1 = 36$ valence electrons
skeleton: four I-Cl single bonds

Step 2 • Build the skeleton, then the outer octets

4 bonds: $4 \times 2 = 8$ • 4 Cl $\times$ 3 lone pairs: $4 \times 6 = 24$
 $8 + 24 = 32$ of 36 placed • every Cl has its octet

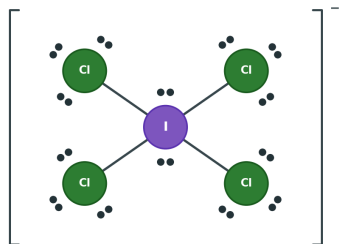
 **Sense check:** Thirty-two of the 36 electrons are placed and every chlorine is full. Four remain, and the only atom left is iodine.

✓ Worked example 2: the expanded center

ICl_4^- : $36 - 8 - 24 = 4$ electrons left
every Cl full · the leftovers belong to iodine

Step 3 · Leftovers go on the central atom

4 leftover $\rightarrow$ 2 lone pairs on I · I: $8 + 4 = 12$ electrons
4 bonds (8) + 2 lone pairs (4) · period 5 $\rightarrow$ expansion allowed



🧠 **Sense check:** Twelve electrons around iodine. Period 5 has the room; a second-period atom never could.

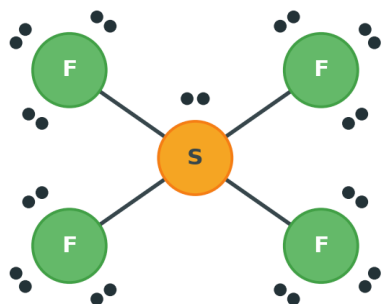
Your turn: sulfur tetrafluoride

SF_4 : $6 + 4(7) = 34$ valence electrons

skeleton: four S-F single bonds · complete the fluorine octets first

$8 + 24 = 32$ · $34 - 32 = 2 \rightarrow 1$ lone pair on S · S: $8 + 2 = 10$

expanded octet on period-3 sulfur · every F keeps its octet



⚠ Where this goes wrong

Expanding a second-period atom. Ten electrons around carbon or nitrogen is never right. Expansion starts in period 3; C, N, O, and F stop at eight.

Dropping the ion's charge. ICl_4^- counted neutral gives $7 + 4(7) = 35$. The $1-$ charge adds the 36th electron; without it the whole structure runs one short.

Pairing the unpairable. NO holds 11 valence electrons. No structure pairs them all; one electron stays single, and the drawing must show it.

Practice 1: sulfur hexafluoride

SF_6 : $6 + 6(7) = 48$ valence electrons · six S-F single bonds
each F carries 3 lone pairs · nothing is left over

How many electrons surround the sulfur atom?

- A. 48
- B. 36
- C. 18
- D. 12

✓ Practice 1 answer: D

S: 6 bonds $\times$ 2 = 12 electrons $\rightarrow$ answer D
period-3 sulfur holds all six bonds · an expanded octet

C added sulfur's own six valence electrons on top of the twelve in its bonds: $12 + 6 = 18$; those six are already inside the bonds. A counted the whole pool: $6 + 6(7) = 48$. B counted the fluorine lone pairs: $6 \times 6 = 36$.

 **Sense check:** Six bonds alone put 12 electrons on sulfur. The 36 lone-pair electrons live on the fluorines, not on the center.



Worked example 3: boron trifluoride

BF_3 : $3 + 3(7) = 24$ · three B-F bonds + nine F lone pairs = 24
all placed · boron sits at six electrons

Tempting fix: move a fluorine lone pair into a B=F double bond and give boron an octet.
Test the fix with formal charges.

✓ Worked example 3: the fix fails

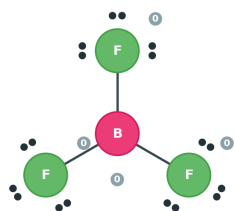
candidate: one B=F double bond · boron reaches 8 · $\text{BF}_3: 3 + 3(7) = 24$

Formal charge on the double-bonded F

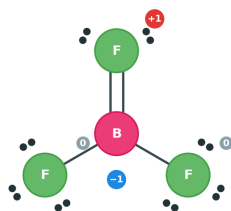
Formal charge on boron

B: $3 - 0 - 4 = -1$ · F: $7 - 4 - 2 = +1$ · all-single structure: every FC = 0

B: $3 - 0 - 3 = 0$ · each F: $7 - 6 - 1 = 0$



A · single bonds, B at six



B · forced B=F double bond

🧠 **Sense check:** Negative on B, positive on F: backward. The all-single structure wins.

Take-home: an incomplete octet can be the best structure

When completing an octet forces bad formal charges, leave the octet incomplete. Boron and beryllium accept fewer than eight; formal charge makes the call.

BF_3 best structure: three single bonds · B at 6 · every FC = 0
forcing $\text{B}=\text{F}$ gives -1 on B and $+1$ on F · rejected

Practice 2: completing ICl_2^-

ICl_2^- : skeleton Cl-I-Cl

given: two I-Cl single bonds · three lone pairs on each Cl · wanted: what completes the structure

A student builds the ion with two single bonds and a full octet on each chlorine, then stops. What completes the structure?

- A. Two lone pairs and one unpaired electron on iodine
- B. Two lone pairs on iodine, giving it an octet
- C. Three lone pairs on iodine, giving it 10 electrons
- D. One lone pair on iodine and one extra pair on each chlorine

✓ Practice 2 answer: C

$$\text{ICl}_2^-: 7 + 2(7) + 1 = 22 \cdot \text{placed: } 2 \text{ bonds (4)} + 6 \text{ Cl lone pairs (12)} = 16 \cdot \text{left: } 22 - 16 = 6$$

6 leftover $\rightarrow$ 3 lone pairs on I $\cdot$ I: $4 + 6 = 10$ electrons $\cdot$ period 5 $\rightarrow$ allowed $\rightarrow$ answer C

A dropped the 1– charge: $7 + 2(7) = 21$ leaves 5, two pairs and a stray electron; the added electron makes 22, an even count that pairs fully. B subtracted the charge instead: $7 + 2(7) - 1 = 20$ leaves 4, two pairs and an octet; a negative charge adds an electron. D placed the right 22 but parked the leftovers on the chlorines; every outer atom already has its octet, so leftovers go on the center.

 **Sense check:** Iodine ends with two bonds and three lone pairs, ten electrons: an expanded octet on a period-5 center.

Check yourself

1. Which periods allow an expanded octet, and which four elements never expand?
2. Count the electrons around sulfur in SF_4 and around iodine in ICl_4^- .

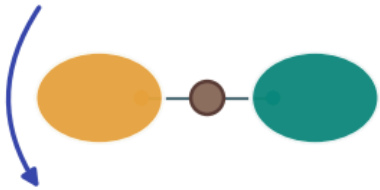
The count of bonds and lone pairs around a central atom, octet or expanded, is exactly what fixes a molecule's three-dimensional shape.

5 · VSEPR & Molecular Shape

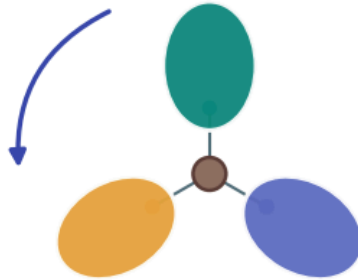
Predict a molecule's electron-pair geometry and its shape by counting the electron domains on the central atom and reading which corners lone pairs take.

What sets a molecule's shape

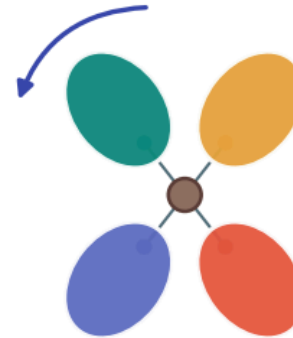
Tied at one knot, balloons crowd, then spread as far apart as they can.



2 balloons: a straight line



3 balloons: a flat triangle



4 balloons: they lift into 3-D

Tie balloons at one knot and they push apart to the roomiest arrangement. Atoms bonded to a central atom spread out the very same way.

Electron domains repel and spread apart

An electron domain is one group of electrons on the central atom: a bond or a lone pair. Like charges repel, so the domains spread as far apart as they can.

**domains push to the farthest-apart arrangement
that arrangement, lone pairs included, sets the molecule's shape**

Counting the domains

Count the domains on the central atom: one for each bonded atom, plus one for each lone pair. A single, double, or triple bond counts as one domain, no matter how many pairs it holds.

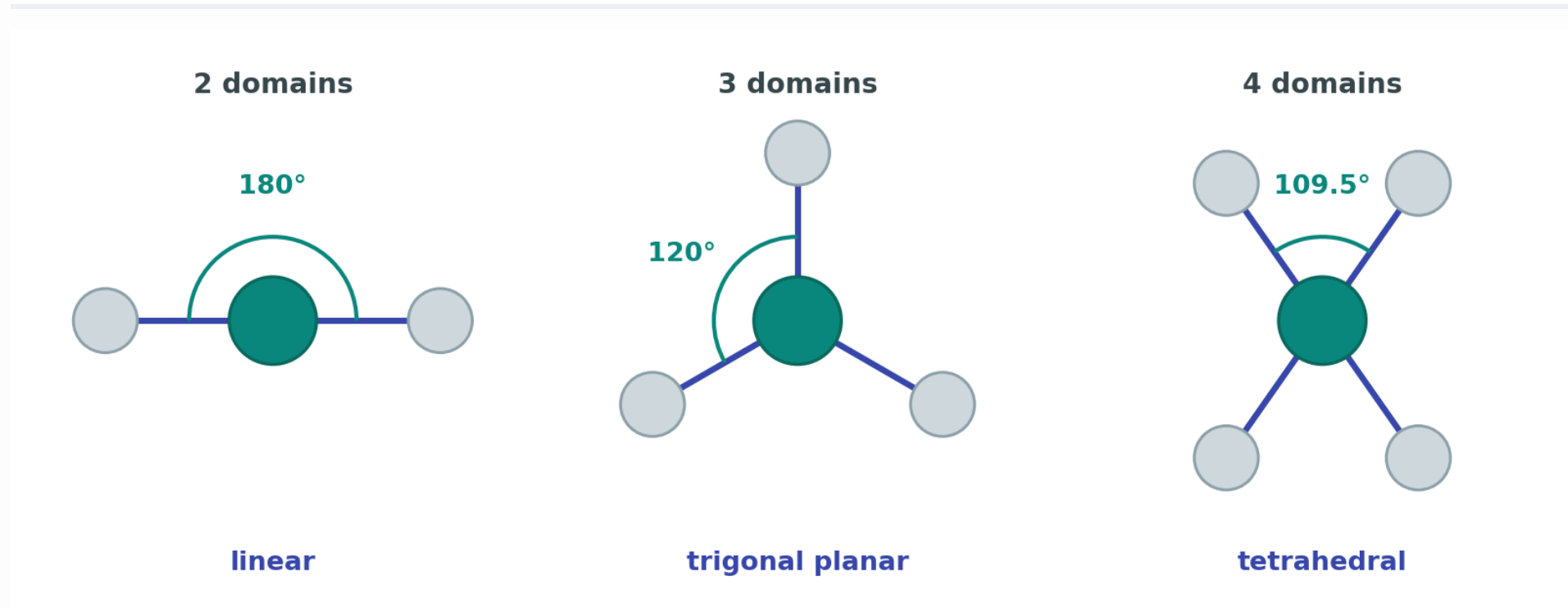
CO₂: C bonded to 2 O by double bonds

2 bonded atoms + 0 lone pairs → 2 domains (each double bond counts once)

NH₃: N bonded to 3 H, plus 1 lone pair

3 bonded atoms + 1 lone pair → 4 domains

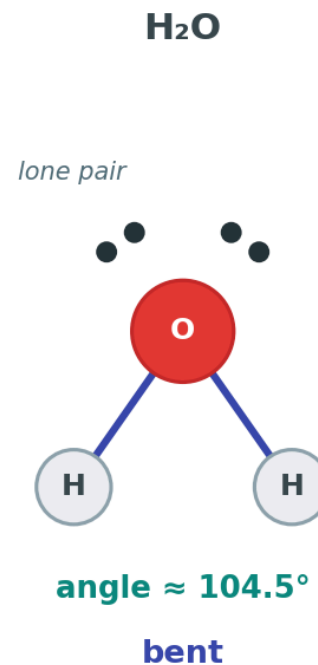
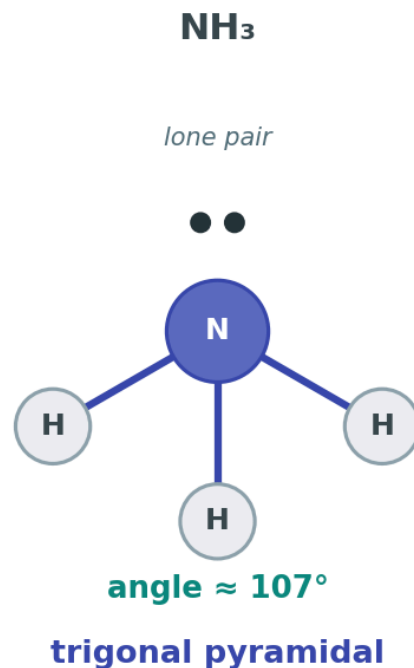
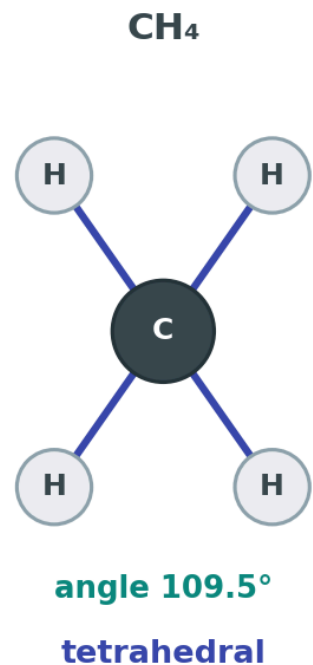
From domain count to electron-pair geometry



The number of domains fixes how they arrange in space. Two domains sit at 180° , three at 120° , four at 109.5° . This spread is the electron-pair geometry.

From geometry to molecular shape

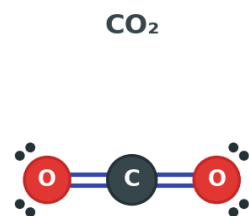
Same electron arrangement, 4 tetrahedral domains: three different shapes



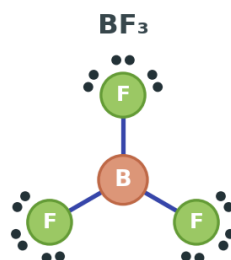
Lone pairs take up domains but hold no atom. The molecular shape names only where the atoms sit. Four tetrahedral domains read as tetrahedral, trigonal pyramidal, or bent as lone pairs replace atoms.

The common shapes side by side

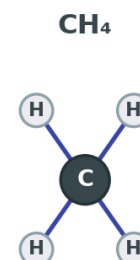
The lone pairs are the dark dots: they claim a domain but hold no atom, so the shape reads only where the atoms sit.



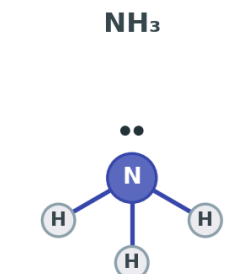
linear
AX₂ · 180°



trigonal planar
AX₃ · 120°



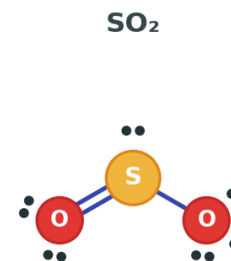
tetrahedral
AX₄ · 109.5°



trigonal pyramidal
AX₃E · ≈107°



bent
AX₂E₂ · ≈104.5°



bent
AX₂E · ≈120°

memory hook: count the domains, then subtract the lone pairs from the name
2 → linear · 3 → trigonal planar, then bent · 4 → tetrahedral, then trigonal pyramidal, then bent



The method

1. **Count the electron domains.** One per bonded atom, one per lone pair; every bond counts once.
2. **Name the electron-pair geometry.** 2 linear, 3 trigonal planar, 4 tetrahedral.
3. **Name the molecular shape.** Keep atom positions; lone-pair corners stay empty.



Worked example 1: methane

CH₄: C bonded to 4 H

central atom: carbon · 4 bonded H · 0 lone pairs on C

Methane's central carbon bonds to four hydrogens with no lone pairs left over. Give its electron-pair geometry and its molecular shape.

✓ Worked example 1: solution

CH_4 : C bonded to 4 H

central atom: carbon · 4 bonded H · 0 lone pairs on C

Step 1 · Count the electron domains

Step 2 · Name the electron-pair geometry

Step 3 · Name the molecular shape

$\text{CH}_4 \rightarrow$ tetrahedral

4 domains · 0 lone pairs · every corner holds an atom · 109.5°

 **Sense check:** No lone pairs, so every corner holds an atom and the shape matches the electron-pair geometry: tetrahedral.

✓ Worked example 1: the row in the table

CH_4 : $X = 4$ bonded H · $E = 0$ lone pairs $\rightarrow \text{AX}_4$
 found: 4 domains · tetrahedral geometry · tetrahedral shape · 109.5°

A = central atom · X = a bonded atom · E = a lone pair

Electron-domain geometry	Molecular shape	AX_nE_m	Bonded atoms (X)	Lone pairs (E)	Example	Bond angle
2 domains linear	linear	AX_2	2	0	CO_2	180°
3 domains trigonal planar	trigonal planar	AX_3	3	0	BF_3	120°
	bent	AX_2E	2	1	SO_2	$\approx 120^\circ$
4 domains tetrahedral	tetrahedral	AX_4	4	0	CH_4	109.5°
	trigonal pyramidal	AX_3E	3	1	NH_3	$\approx 107^\circ$
	bent	AX_2E_2	2	2	H_2O	$\approx 104.5^\circ$

🧠 **Sense check:** Four atoms and no lone pairs: the AX_4 row. The geometry and the shape share one name. ✓



Worked example 2: carbon dioxide

O=C=O: C bonded to 2 O by double bonds
central atom: carbon · 2 bonded O · 0 lone pairs on C

Carbon dioxide holds two carbon-oxygen double bonds. A common first attempt: two double bonds make four domains. Find the geometry and the shape.

✓ Worked example 2: solution

$\text{O}=\text{C}=\text{O}$: C bonded to 2 O by double bonds
central atom: carbon · 2 bonded O · 0 lone pairs on C

A common first attempt

2 double bonds = 4 domains?

counting each double bond as two: $2 \times 2 = 4$ × · a bond is one domain however many pairs it holds

Step 1 · Count the electron domains

Step 2 · Name the electron-pair geometry

Step 3 · Name the molecular shape

$\text{CO}_2 \rightarrow \text{linear}$

2 domains · 0 lone pairs · both corners hold an atom · 180°

 **Sense check:** Each double bond is one domain, so carbon has two. Two domains can only point 180° apart: a linear molecule.

✓ Worked example 2: the row in the table

CO_2 : $X = 2$ bonded O · $E = 0$ lone pairs $\rightarrow \text{AX}_2$
found: 2 domains · linear geometry · linear shape · 180°

A = central atom · X = a bonded atom · E = a lone pair

Electron-domain geometry	Molecular shape	AX_nE_m	Bonded atoms (X)	Lone pairs (E)	Example	Bond angle
2 domains linear	linear	AX_2	2	0	CO_2	180°
3 domains trigonal planar	trigonal planar	AX_3	3	0	BF_3	120°
	bent	AX_2E	2	1	SO_2	$\approx 120^\circ$
4 domains tetrahedral	tetrahedral	AX_4	4	0	CH_4	109.5°
	trigonal pyramidal	AX_3E	3	1	NH_3	$\approx 107^\circ$
	bent	AX_2E_2	2	2	H_2O	$\approx 104.5^\circ$

🧠 **Sense check:** Two atoms and no lone pairs: the AX_2 row. Each double bond counts once, so the row stays AX_2 . ✓

Take-home: a bond is one domain

$\text{O}=\text{C}=\text{O} \rightarrow 2 \text{ domains} \rightarrow \text{linear } (180^\circ)$

each double bond counts once: $1 + 1 = 2 \checkmark$

counting each double bond twice $\rightarrow 4 \text{ domains} \rightarrow \text{a tetrahedral guess}$

$2 \times 2 = 4 \times$ · the extra domains would bend a straight molecule

Bond order does not change the domain count. A single, double, or triple bond is one region of electrons, so it claims one domain.

Your turn: ammonia

NH₃: N bonded to 3 H, with 1 lone pair on N
 central atom: nitrogen · 3 bonded H · 1 lone pair

step	count	result
1 · domains	3 bonded + 1 lone pair	_____ domains
2 · electron-pair geometry	from 4 domains	_____
3 · molecular shape	3 atoms, 1 corner a lone pair	_____

Fill the three results.

NH₃ → trigonal pyramidal

3 + 1 = 4 domains · tetrahedral geometry · 1 corner empty · angle ≈ 107°

⚠ Where this goes wrong

Reporting the electron-pair geometry as the shape. Water has four domains, a tetrahedral geometry. But two corners are lone pairs. The shape names only the atoms: bent, not tetrahedral.

Ignoring the lone pairs. Arrange only NH_3 's three bonded atoms and you get a flat trigonal planar. The lone pair takes a corner too, pressing the atoms into a trigonal pyramid.

Counting a double bond as two domains. SO_2 has two bonded oxygens plus one lone pair. Count the $\text{S}=\text{O}$ double bond as two and you reach four domains. Each bond is one domain: $2 + 1 = 3$ domains, a bent molecule.



Practice 1

NF_3 : N bonded to 3 F, with 1 lone pair on N
central atom: nitrogen · 3 bonded F · 1 lone pair

Nitrogen trifluoride has three bonded fluorines and one lone pair on nitrogen. What is its molecular shape?

- A. Trigonal pyramidal: three bonded atoms with the lone pair pressing them down
- B. Tetrahedral: the four electron domains arrange as a tetrahedron
- C. Trigonal planar: the three fluorines spread evenly around nitrogen
- D. Bent: the lone pair leaves only a bent arrangement

✓ Practice 1 answer: A

$\text{NF}_3 \rightarrow$ trigonal pyramidal $\rightarrow$ answer A

3 bonded F + 1 lone pair = 4 domains · tetrahedral geometry · 1 corner empty

B named the electron-pair geometry, not the shape: the lone pair takes a corner, so the atoms are not tetrahedral. C ignored the lone pair; three bonded atoms alone would be trigonal planar, but the fourth corner is filled. D miscounts: four domains, not three, so the base is a tetrahedron, not a triangle.

 **Sense check:** Four domains, one a lone pair: the three fluorines press down into a pyramid. Trigonal pyramidal.



Worked example 3: sulfur dioxide

O=S-O: S bonded to 2 O, with 1 lone pair on S

central atom: sulfur · 2 bonded O · 1 lone pair · one bond is double

Sulfur dioxide has two bonded oxygens, one of them a double bond, and a lone pair on sulfur. Give its electron-pair geometry and its molecular shape.

✓ Worked example 3: solution

$\text{O}=\text{S}-\text{O}$: S bonded to 2 O, with 1 lone pair on S
central atom: sulfur · 2 bonded O · 1 lone pair · one bond is double

Step 1 · Count the electron domains

Step 2 · Name the electron-pair geometry

Step 3 · Name the molecular shape

$\text{SO}_2 \rightarrow$ bent

3 domains · trigonal planar geometry · 1 corner a lone pair · angle just under 120°

 **Sense check:** Three domains, one a lone pair, leave two oxygens in a bent line. The lone pair squeezes the angle a little below 120° .

✓ Worked example 3: the row in the table

SO_2 : X = 2 bonded O · E = 1 lone pair $\rightarrow \text{AX}_2\text{E}$

found: 3 domains · trigonal planar geometry · bent shape · just under 120°

A = central atom · X = a bonded atom · E = a lone pair

Electron-domain geometry	Molecular shape	AX_nE_m	Bonded atoms (X)	Lone pairs (E)	Example	Bond angle
2 domains linear	linear	AX_2	2	0	CO_2	180°
3 domains trigonal planar	trigonal planar	AX_3	3	0	BF_3	120°
	bent	AX_2E	2	1	SO_2	$\approx 120^\circ$
4 domains tetrahedral	tetrahedral	AX_4	4	0	CH_4	109.5°
	trigonal pyramidal	AX_3E	3	1	NH_3	$\approx 107^\circ$
	bent	AX_2E_2	2	2	H_2O	$\approx 104.5^\circ$

🧠 **Sense check:** Two atoms and one lone pair: the AX_2E row, inside the trigonal planar group. ✓

Summary: domains, geometry, shape

VSEPR SUMMARY: domains, geometry, shape

A = central atom · X = a bonded atom · E = a lone pair

Electron-domain geometry	Molecular shape	AX_nE_m	Bonded atoms (X)	Lone pairs (E)	Example	Bond angle
2 domains linear	linear	AX_2	2	0	CO_2	180°
3 domains trigonal planar	trigonal planar	AX_3	3	0	BF_3	120°
	bent	AX_2E	2	1	SO_2	$\approx 120^\circ$
4 domains tetrahedral	tetrahedral	AX_4	4	0	CH_4	109.5°
	trigonal pyramidal	AX_3E	3	1	NH_3	$\approx 107^\circ$
	bent	AX_2E_2	2	2	H_2O	$\approx 104.5^\circ$

Five and six domains (trigonal bipyramidal, octahedral) extend this table; the common molecules use two, three, or four.

Count the bonded atoms (X) and lone pairs (E), then read the row across.



Worked example 4: water

H-O-H: O bonded to 2 H, with 2 lone pairs on O

central atom: oxygen · 2 bonded H · 2 lone pairs on O

Water's oxygen bonds to two hydrogens and keeps two lone pairs. Count the domains, then read its row for the electron-pair geometry, the shape, and the angle.

✓ Worked example 4: solution

H-O-H: O bonded to 2 H, with 2 lone pairs on O
central atom: oxygen · 2 bonded H · 2 lone pairs on O

Step 1 · Count the electron domains

Step 2 · Name the electron-pair geometry

Step 3 · Name the molecular shape



4 domains · tetrahedral geometry · 2 corners hold lone pairs · angle $\approx 104.5^\circ$

 **Sense check:** Two of the four corners hold lone pairs. The two hydrogens are left in a bent line, near 104.5° .

✓ Worked example 4: the row in the table

H_2O : $X = 2$ bonded H · $E = 2$ lone pairs $\rightarrow \text{AX}_2\text{E}_2$
found: 4 domains · tetrahedral geometry · bent shape · about 104.5°

A = central atom · X = a bonded atom · E = a lone pair

Electron-domain geometry	Molecular shape	AX_nE_m	Bonded atoms (X)	Lone pairs (E)	Example	Bond angle
2 domains linear	linear	AX_2	2	0	CO_2	180°
3 domains trigonal planar	trigonal planar	AX_3	3	0	BF_3	120°
	bent	AX_2E	2	1	SO_2	$\approx 120^\circ$
4 domains tetrahedral	tetrahedral	AX_4	4	0	CH_4	109.5°
	trigonal pyramidal	AX_3E	3	1	NH_3	$\approx 107^\circ$
	bent	AX_2E_2	2	2	H_2O	$\approx 104.5^\circ$

 **Sense check:** Two atoms and two lone pairs: the AX_2E_2 row. Same shape name as SO_2 , a different row and a smaller angle. ✓



Practice 2

H₂Se: Se bonded to 2 H

a student draws two Se-H bonds, no lone pairs on Se, and names the shape linear

Evaluate the work.

- A. Correct: two bonded atoms give two domains, 180° apart
- B. Wrong: selenium keeps two lone pairs, so four domains make the shape tetrahedral
- C. Wrong: selenium keeps two lone pairs, so four domains make the shape bent
- D. Wrong: selenium keeps one lone pair, so three domains make the shape bent, near 120°

✓ Practice 2 answer: C

H_2Se : $2(1) + 6 = 8$ valence electrons · $8 - 4$ in the two bonds = 4 left → 2 lone pairs on Se

$2 + 2 = 4$ domains · tetrahedral geometry · 2 corners hold lone pairs · bent shape → answer C

A dropped the lone pairs: only 4 of the 8 electrons sit in the two bonds, and the other 4 stay on selenium as two more domains. B named the electron-pair geometry; two of the four corners hold lone pairs, so the atoms are bent, not tetrahedral. D miscounted the leftovers as one pair: $8 - 4 = 4$ electrons make two pairs, four domains, not three.

 **Sense check:** Same count as water: two bonds, two lone pairs, four domains. The drawing must show the lone pairs before the shape can be read.



Check yourself

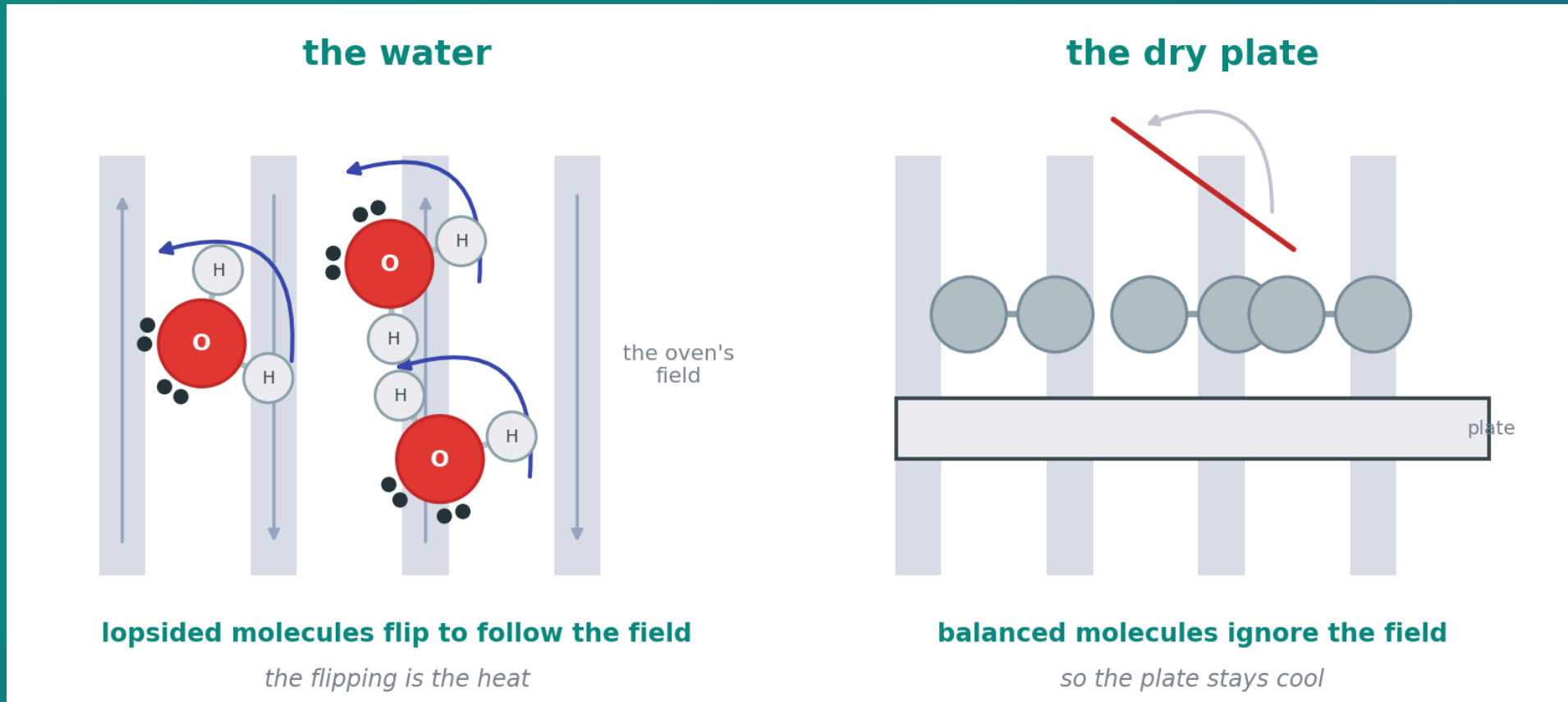
1. PH_3 has a phosphorus bonded to three hydrogens and one lone pair. Count the domains, name the electron-pair geometry, then the shape.
2. Both CO_2 and SO_2 have two bonded oxygens. Why is one linear and the other bent?

Shape sets polarity. Two equal bond dipoles cancel when they point exactly opposite, as in linear CO_2 . A bent or pyramidal shape leaves them pointing partly the same way, so the molecule is polar. Combining these bond dipoles with the molecular shape shows whether the whole molecule is polar.

6 • Molecular Polarity

Decide whether a whole molecule is polar by combining its bond dipoles with its shape: symmetric shapes with identical outer atoms cancel the dipoles to nonpolar, while lopsided shapes or a central lone pair leave a net dipole.

Why the plate stays cool



A microwave heats the soup, not the dry plate under it. Water molecules are lopsided, so the oven's field keeps twisting them. That twisting is the heat.

The shape decides, not the bonds

Each polar bond carries a dipole toward its more electronegative atom. The molecule is polar only when these dipoles do not cancel. The shape decides whether they cancel.

CO₂: polar bonds, the dipoles cancel
nonpolar molecule

H₂O: polar bonds, the dipoles add
polar molecule

Every polar bond is an arrow

a bond dipole

points to the more electronegative atom

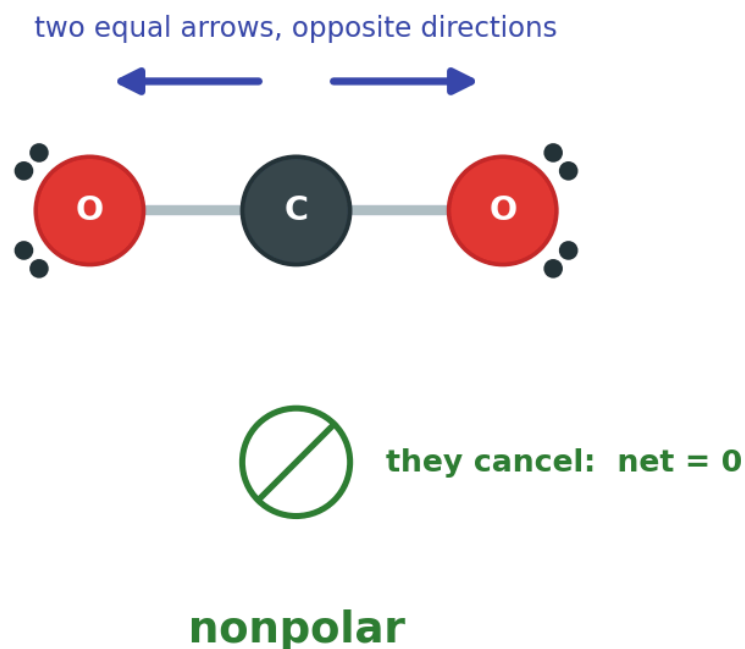


the shared electrons sit closer to chlorine

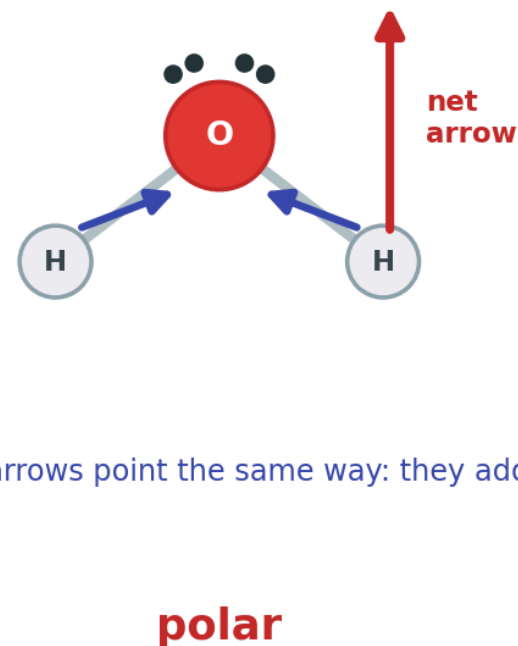
Draw each polar bond as a dipole arrow. It points to the more electronegative atom, the end that pulls the shared electrons closer. A bigger electronegativity difference means a stronger pull.

When arrows cancel, and when they add

CO₂ · linear



H₂O · bent



Identical arrows arranged evenly around the center cancel, and the molecule is nonpolar. A lopsided shape, or a central lone pair, leaves a net arrow, and the molecule is polar.



The method

1. **Recall the shape** from VSEPR.
2. **Draw the bond dipoles:** one arrow per bond, toward the more electronegative atom.
3. **Add the arrows.** Identical arrows in a symmetric shape cancel; a leftover arrow means polar.



Worked example 1: carbon dioxide



two C=O bonds · wanted: polar or nonpolar?

Carbon dioxide has two polar C=O bonds.

A common first answer: polar bonds, so the molecule is polar. Test it against the shape.

✓ Worked example 1: solution



two C=O bonds · wanted: polar or nonpolar?

Step 1 · Recall the shape

Step 2 · Draw the bond dipoles

$\Delta\text{EN}(\text{C}=\text{O}): 3.44 - 2.55 = \mathbf{0.89}$ · each bond polar, arrow toward O

Step 3 · Add the arrows

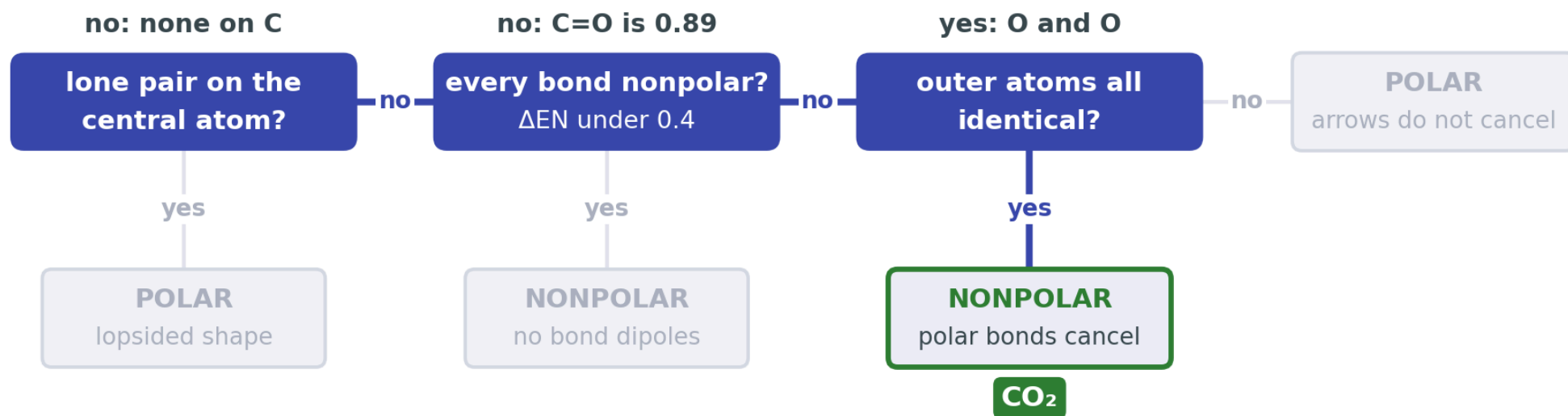
0.89 toward one O – 0.89 toward the other O = **0 net**

 **Sense check:** The C=O bonds are polar, yet CO₂ is nonpolar. Its linear shape aims the two equal arrows in opposite directions.

✓ Worked example 1: the route on the flowchart

CO_2 : $\text{O}=\text{C}=\text{O}$, linear

no lone pair on C · C=O polar (0.89) · two identical O atoms · found: nonpolar



🧠 **Sense check:** No lone pair on carbon, polar bonds, identical outer atoms: the two arrows cancel.

Take-home: polar bonds do not make a polar molecule

CO₂: two polar C=O bonds, linear
symmetric: the two arrows cancel → nonpolar

H₂O: two polar O-H bonds, bent
lopsided: the two arrows add → polar

Both molecules have polar bonds. The linear shape cancels them; the bent shape does not. Polar bonds alone are not enough. The shape decides.



Worked example 2: water

H₂O: two O-H bonds

two lone pairs on oxygen · wanted: polar or nonpolar?

Water has two polar O-H bonds and two lone pairs on its oxygen.

Apply the three steps.

✓ Worked example 2: solution

H₂O: two O-H bonds

two lone pairs on oxygen · wanted: polar or nonpolar?

Step 1 · Recall the shape

Step 2 · Draw the bond dipoles

$\Delta\text{EN}(\text{O}-\text{H}): 3.44 - 2.20 = 1.24$ · each bond polar, arrow toward O

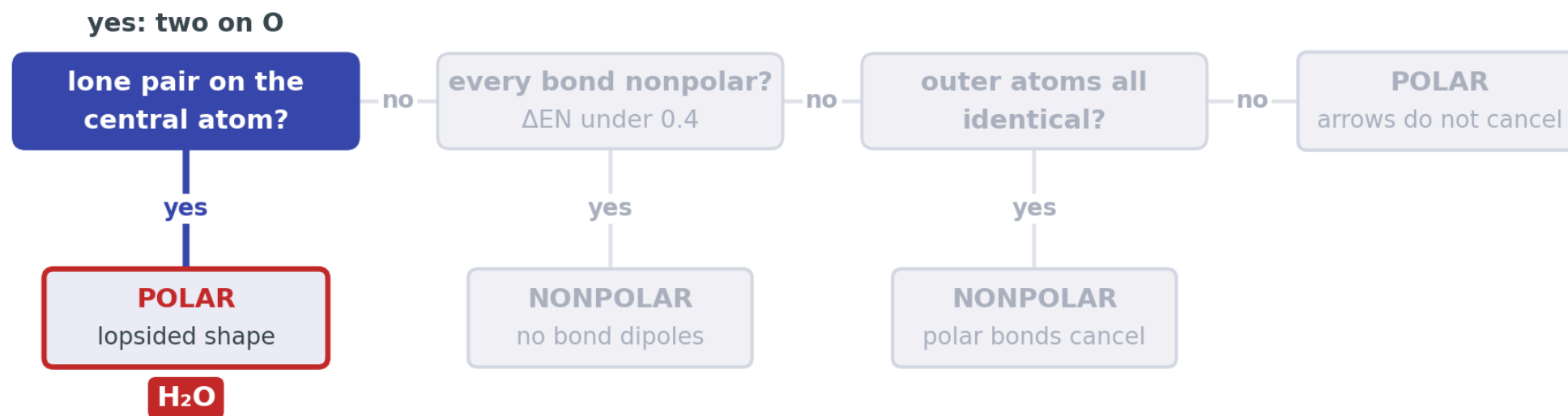
Step 3 · Add the arrows

arrow toward O + arrow toward O → **one net arrow through O**

 **Sense check:** Both molecules have polar bonds, opposite results. Water's bent shape lets the two arrows add, and the lone pairs keep it from ever being symmetric.

✓ Worked example 2: the route on the flowchart

H_2O : bent, two lone pairs on O
two lone pairs on O · O-H polar (1.24) · found: polar



Sense check: The first question settles it: lone pairs on the central oxygen leave the molecule lopsided. ✓

Your turn: ammonia

NH_3 : three N-H bonds, one lone pair on nitrogen

$\Delta\text{EN}(\text{N-H}) = 3.04 - 2.20 = 0.84 \rightarrow$ each bond polar

shape: — · three arrows toward N · they — $\rightarrow$ NH_3 is —

Recall the shape, draw the three arrows, then decide whether they cancel.

shape: trigonal pyramidal · three arrows toward N, tilted to one side · they **do not cancel**

$\rightarrow$ NH_3 is **polar**

The lone pair pushes the three N-H bonds to one side, so their arrows cannot balance. A net arrow points through the nitrogen.

⚠ Where this goes wrong

Polar bonds, so a polar molecule. CO_2 has two polar $\text{C}=\text{O}$ bonds, and the leap to "so CO_2 is polar" skips the shape. Linear and symmetric, its two arrows cancel: nonpolar. The bonds do not decide; the shape does.

Calling a symmetric shape lopsided. SO_3 is trigonal planar with three identical $\text{S}=\text{O}$ bonds and no lone pair on sulfur. Claiming the arrows "add up" ignores that three equal arrows 120° apart cancel exactly. Identical outer atoms in an even arrangement balance.

Nonpolar for the wrong reason. CCl_4 is nonpolar, but not because its bonds are nonpolar. Each $\text{C}-\text{Cl}$ bond is polar. The molecule is nonpolar because the tetrahedral shape cancels the four arrows.

Missing the lone pair. Treating NH_3 as a flat, even molecule calls it nonpolar. The lone pair on nitrogen tilts the three $\text{N}-\text{H}$ arrows to one side, and they no longer cancel: polar.



Practice 1

BCl_3 : three B-Cl bonds, trigonal planar, no lone pair on boron

$$\Delta\text{EN}(\text{B-Cl}) = 3.16 - 2.04 = 1.12 \rightarrow \text{each bond polar}$$

Boron trichloride has three identical polar bonds in a flat triangle. Polar or nonpolar, and why?

- A. Polar: the trigonal planar shape places the three arrows asymmetrically, so they add up
- B. Polar: it contains polar B-Cl bonds, so the whole molecule must be polar
- C. Nonpolar: the B-Cl bonds are polar, but the trigonal planar shape is symmetric, so the three arrows cancel
- D. Nonpolar: none of its B-Cl bonds are polar in the first place

✓ Practice 1 answer: C

$\text{BCl}_3 \rightarrow \text{nonpolar} \rightarrow \text{answer C}$

three polar B-Cl arrows, 120° apart, cancel $\rightarrow$ no net arrow

trigonal planar, symmetric · three arrows toward Cl, 120° apart $\rightarrow$ **cancel: nonpolar**

B took polar bonds as proof of a polar molecule and skipped the shape: the symmetric triangle cancels the three equal arrows. A called the symmetric shape lopsided, but three identical arrows 120° apart balance exactly. D denied the bonds are polar. $\Delta\text{EN}(\text{B-Cl}) = 3.16 - 2.04 = 1.12$, so each bond is polar, and the molecule is nonpolar because of the shape, not because the bonds are nonpolar.

 **Sense check:** Three equal arrows spread evenly around the boron sum to zero. Symmetric polar bonds $\rightarrow$ nonpolar. ✓



Worked example 3: two tetrahedral molecules

CCl_4 and CHCl_3 : both tetrahedral, carbon at the center
four C-Cl bonds · CHCl_3 swaps one Cl for H · wanted: each polar or nonpolar?

Two molecules with the same tetrahedral shape. CCl_4 has four C-Cl bonds; CHCl_3 replaces one chlorine with a hydrogen.

Judge each: polar or nonpolar?

✓ Worked example 3: shape and dipoles

CCl_4 and CHCl_3 : both tetrahedral, carbon at the center
wanted: each molecule polar or nonpolar?

Step 1 · Recall the shape

Step 2 · Draw the bond dipoles

$$\Delta\text{EN}(\text{C-Cl}): 3.16 - 2.55 = \mathbf{0.61} \quad \cdot \quad \Delta\text{EN}(\text{C-H}): 2.55 - 2.20 = \mathbf{0.35}$$

· C-Cl polar · C-H under 0.4, nearly nonpolar

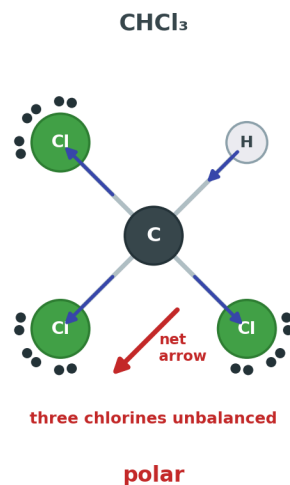
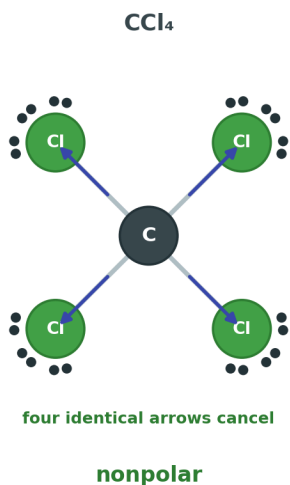
 **Sense check:** The C-Cl bonds are polar. Whether the molecule is polar now depends only on how the four arrows are arranged.

✓ Worked example 3: adding the arrows

same tetrahedral shape, one outer atom swapped

$\text{CCl}_4 \rightarrow \text{nonpolar} \cdot \text{CHCl}_3 \rightarrow \text{polar}$

Step 3 • Add the arrows

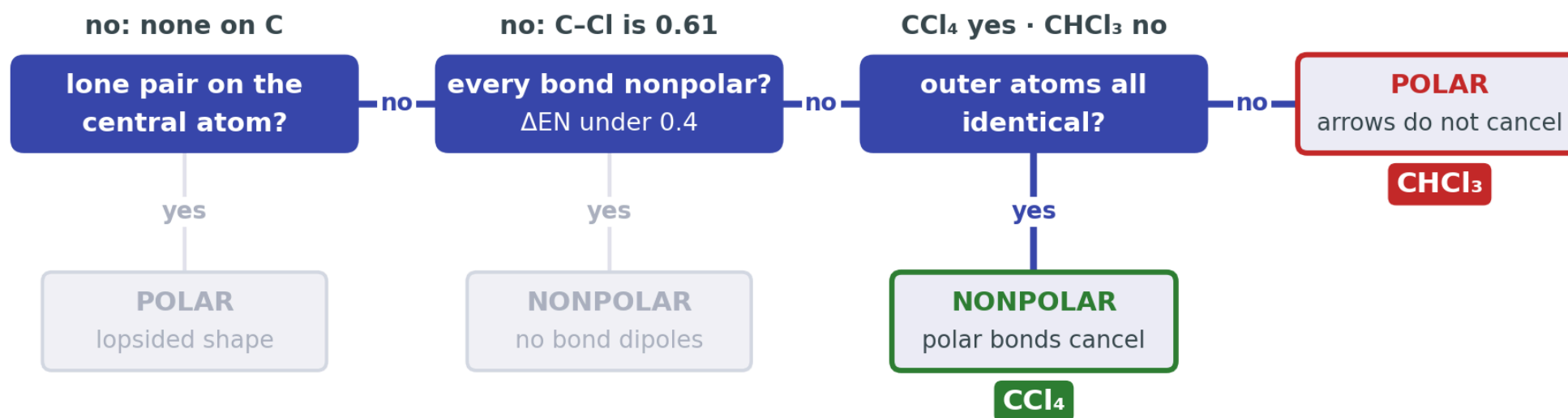


🧠 **Sense check:** CCl_4 's four identical arrows cancel; swapping one chlorine for hydrogen leaves the three chlorine arrows unbalanced, so a net arrow remains: polar. ✓

✓ Worked example 3: the route on the flowchart

CCl_4 and CHCl_3 : tetrahedral, no lone pair on C

no lone pair on C · C-Cl polar (0.61) · CCl_4 : four identical Cl → nonpolar · CHCl_3 : H and Cl mixed → polar



🧠 **Sense check:** Same shape, same first two answers. The outer atoms alone split the two molecules.



Practice 2



wanted: every nonpolar molecule in the set

Which of these molecules are nonpolar?

- A. SiF_4 , OCS and CF_4
- B. SiF_4 and CF_4
- C. SiF_4 , PCl_3 and CF_4
- D. None of them: every one contains polar bonds

✓ Practice 2 answer: B

nonpolar: SiF_4 and $\text{CF}_4 \rightarrow$ answer B

SiF_4 , CF_4 tetrahedral, four identical arrows cancel · PCl_3 trigonal pyramidal, lone pair $\rightarrow$ polar · OCS linear, O and S pull unequally $\rightarrow$ polar

A read OCS as symmetric like CO_2 ; linear cancels only when both outer atoms are identical, and $\text{C}=\text{O}$ out-pulls $\text{C}=\text{S}$. C missed the lone pair on phosphorus and treated PCl_3 as a flat triangle; three bonds plus one lone pair make it pyramidal and polar. D took polar bonds as proof of a polar molecule; SiF_4 and CF_4 have polar bonds and cancel by shape.

 **Sense check:** Nonpolar needs both: a symmetric shape and identical outer atoms. ✓



Check yourself

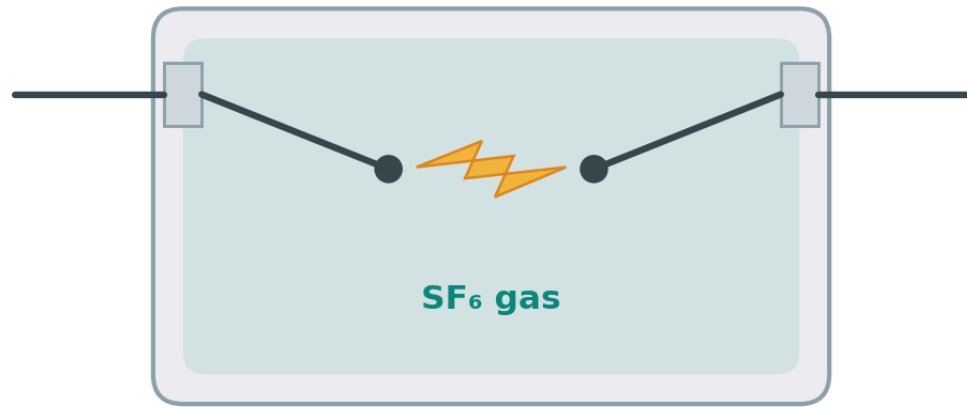
1. BeCl_2 has two polar Be-Cl bonds in a linear shape, with no lone pair on beryllium. Polar or nonpolar? Give the reason.
2. A molecule is built from polar bonds yet turns out nonpolar. What must be true about its shape and its outer atoms?

Polar or nonpolar is now a two-step call: name the shape, then add the arrows. And the logic does not stop at four domains. Central atoms from period 3 on can hold five or six electron groups, and the same symmetry test decides their polarity too.

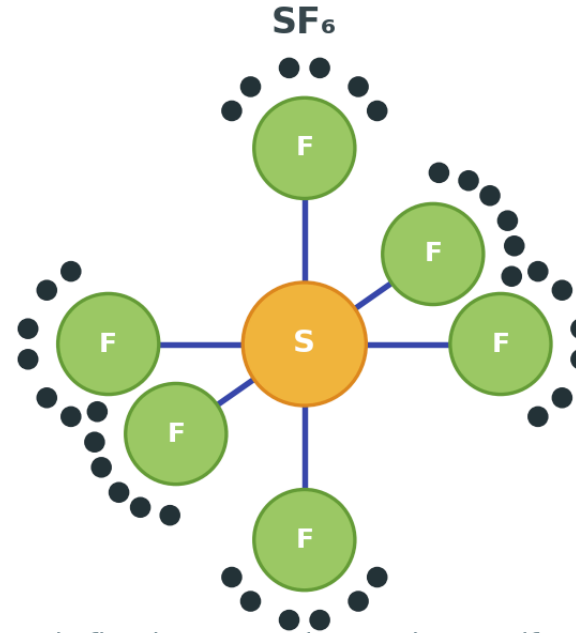
7 • Five & Six Electron Domains

Predict the shapes built on five or six electron domains by seating lone pairs in the roomiest corners, and decide whether the finished molecule is polar.

Six neighbors on one atom



a high-voltage breaker opens under load;
the spark is smothered in SF₆ gas



six fluorines seated around one sulfur

High-voltage breakers quench their sparks in sulfur hexafluoride gas. Six fluorines crowd around one sulfur atom. Four corners cannot seat six neighbors.

Four domains: the shapes you know

CH₄ 4 + 0 → tetrahedral · NH₃ 3 + 1 → trigonal pyramidal · H₂O 2 + 2 → bent
4 domains each · tetrahedral geometry each time · the lone pairs change the shape, not the geometry

Electron domains spread as far apart as they can. Count the domains, name the geometry, then read only the atoms. That routine names every four-domain shape. It works unchanged for five and six.

Five and six domains

Electron domains spread as far apart as they can. That rule does not stop at four. Central atoms from period 3 on can hold five or six domains, an expanded octet.

5 domains → trigonal bipyramidal

3 equatorial corners + 2 axial corners

6 domains → octahedral

6 identical corners · every neighbor at 90°

Trigonal bipyramidal: two kinds of corners

Five domains make two kinds of corners. Three equatorial corners ring the middle; two axial corners cap the ends.

equatorial: 120° from each other · 90° from the axial corners

axial: 180° from each other · 90° from all 3 equatorial corners

Lone pairs take equatorial corners

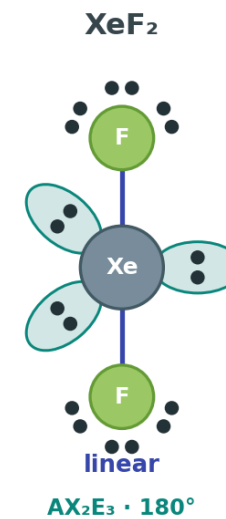
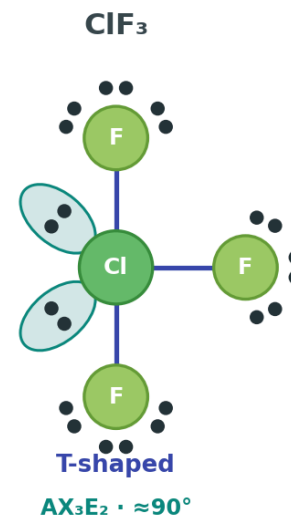
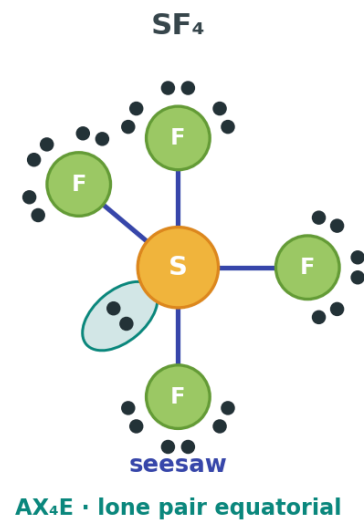
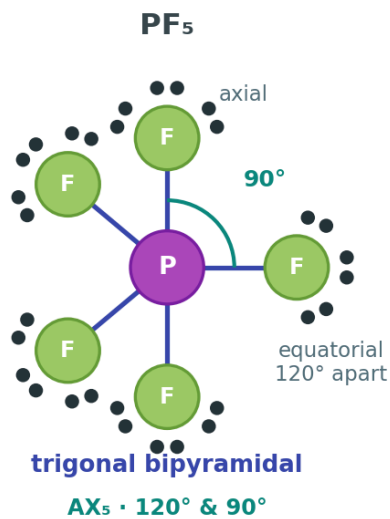
A lone pair is held by one atom, so it spreads wider than a bonding pair. Wide pairs crowd less in an equatorial corner: two neighbors at 90° , not three.

axial lone pair: 3 close contacts at 90° ✗

equatorial lone pair: 2 close contacts at 90° ✓

The five-domain family

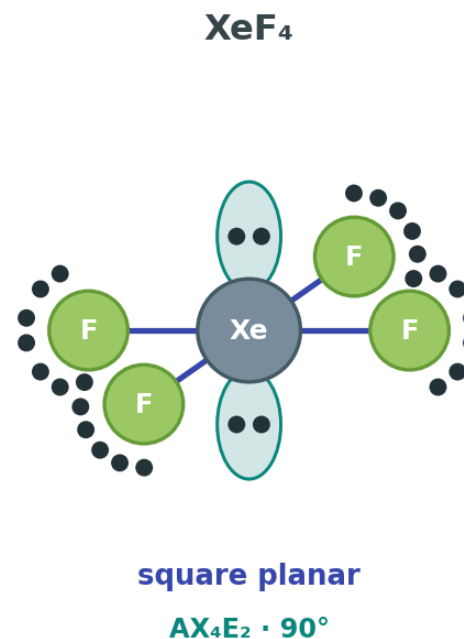
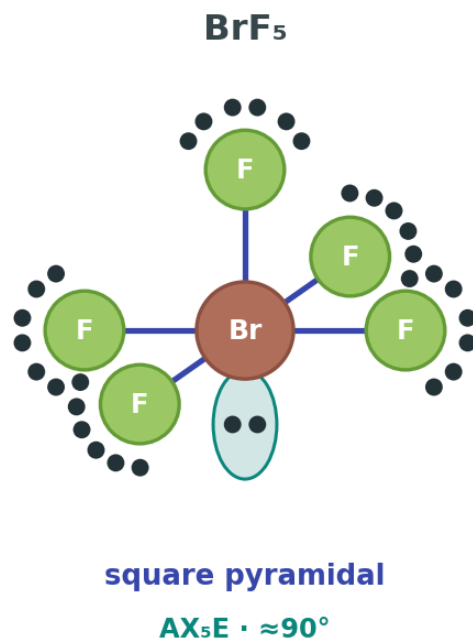
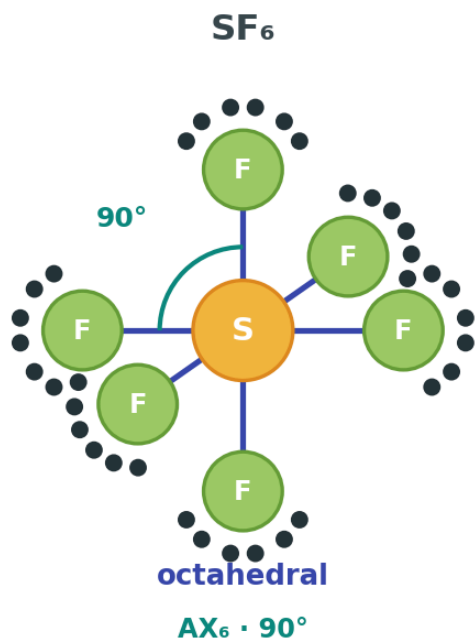
Five domains: three equatorial corners and two axial. The shaded corners hold lone pairs; they fill equatorial spots first.



Lone pairs replace equatorial atoms one at a time. Five domains read as trigonal bipyramidal, seesaw, T-shaped, or linear as the shaded corners fill with lone pairs.

The six-domain family

Six domains: every neighbor at 90°. One lone pair leaves a square pyramid; two sit opposite, leaving a square plane.

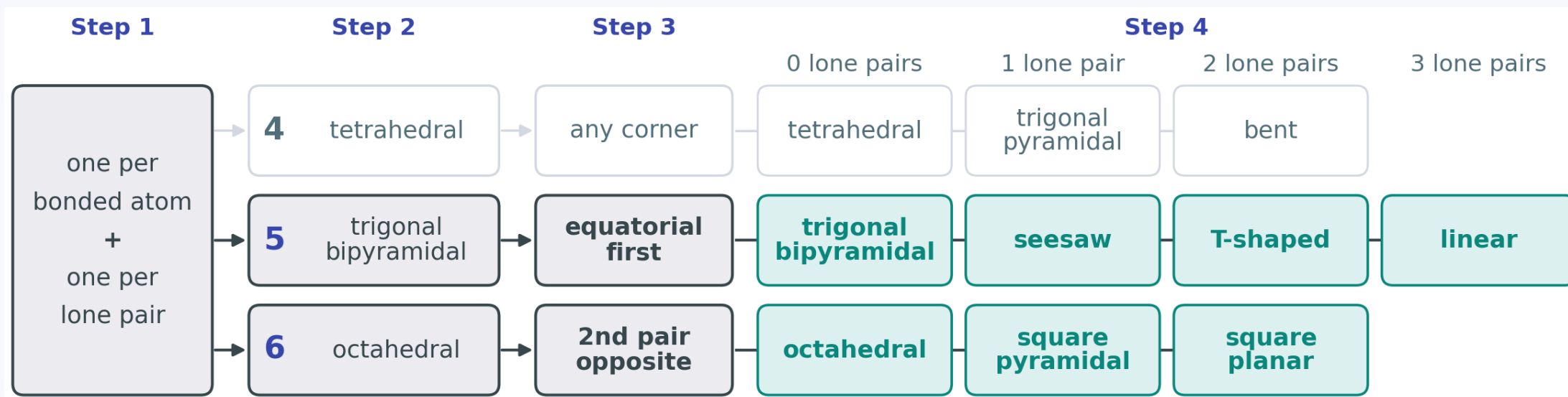


An octahedron's six corners are identical, every neighbor at 90°. One lone pair leaves a square pyramid. A second lone pair sits opposite the first, leaving a square plane.



The method

1. **Count the domains.** One per bonded atom, one per lone pair.
2. **Name the geometry.** 5 trigonal bipyramidal, 6 octahedral.
3. **Seat the lone pairs.** Equatorial first; opposite each other in an octahedron.
4. **Name the shape.** Atom positions only.





Worked example 1: phosphorus pentafluoride

PF_5 : P bonded to 5 F

central atom: phosphorus · 5 bonded F · 0 lone pairs on P

Phosphorus bonds to five fluorines with no lone pairs left over. Give the electron-pair geometry, the molecular shape, and the bond angles.

✓ Worked example 1: solution

PF_5 : P bonded to 5 F

central atom: phosphorus · 5 bonded F · 0 lone pairs on P

Step 1 · Count the domains

Step 2 · Name the geometry

Step 3 · Seat the lone pairs

Step 4 · Name the shape

$\text{PF}_5 \rightarrow$ trigonal bipyramidal

5 domains · 0 lone pairs · 120° equatorial · 90° axial-to-equatorial · 180° axial-to-axial

 **Sense check:** No lone pairs, so the shape is the electron-pair geometry itself. Two kinds of corners give two angles between neighboring bonds, 120° and 90° , plus 180° straight across the axis.





Worked example 2: sulfur tetrafluoride

SF_4 : S bonded to 4 F, with 1 lone pair on S
central atom: sulfur · 4 bonded F · 1 lone pair

Sulfur holds four bonded fluorines and one lone pair. A common first attempt seats the lone pair in an axial corner. Find the molecular shape.

✓ Worked example 2: solution

SF_4 : S bonded to 4 F, with 1 lone pair on S

Step 1 · Count the domains

Step 2 · Name the geometry

A common first attempt

lone pair axial $\rightarrow$ 3 close contacts at $90^\circ \times$

lone pair equatorial $\rightarrow$ 2 close contacts at $90^\circ \checkmark$

Step 3 · Seat the lone pairs

Step 4 · Name the shape

$\text{SF}_4 \rightarrow$ seesaw

5 domains · 1 equatorial lone pair · 2 axial F + 2 equatorial F

🧠 **Sense check:** One equatorial corner holds the lone pair. Its push squeezes the seesaw's angles slightly.



Take-home: lone pairs sit equatorial

equatorial lone pair: 2 neighbors at 90° ✓
axial lone pair: 3 neighbors at 90° ✗ · more crowding

A lone pair spreads wider than a bonding pair. In a trigonal bipyramid every lone pair takes an equatorial corner, and the atoms fit around that choice.



Guided example: krypton difluoride



only the formula · no bonds or lone pairs drawn yet

Krypton is a noble gas, yet it bonds to two fluorines. Find the lone pairs on krypton, then name the molecular shape.

✓ Guided example: the lone pairs

KrF_2 : $8 + 2(7) = 22$ valence electrons

F-Kr-F: 2 bonds use $2 \times 2 = 4$ · each F takes 6 more: $2 \times 6 = 12$ · $22 - 4 - 12 = 6$ left on Kr

Step 1 · Count the domains

KrF_2 : 2 bonded F + 3 lone pairs = 5 domains

Kr holds $4 + 6 = 10$ electrons · period 4, so the octet may expand

 **Sense check:** Stopping at eight would leave 2 electrons unplaced. All 22 must be placed, so krypton carries five domains.

✓ Guided example: the shape

KrF_2 : 2 bonded F + 3 lone pairs on Kr = 5 domains
 from the count: $22 - 4 - 12 = 6$ left $\rightarrow$ 3 lone pairs

Step 2 · Name the geometry

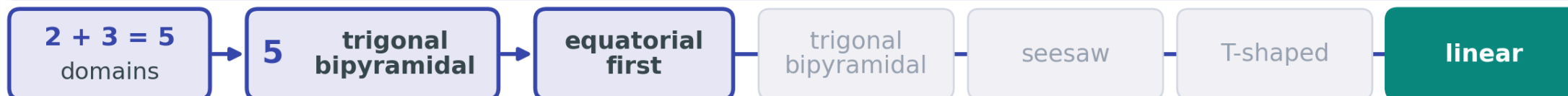
Step 3 · Seat the lone pairs

Step 4 · Name the shape

$\text{KrF}_2 \rightarrow$ linear

5 domains · 3 equatorial lone pairs · 2 axial F · $\text{F-Kr-F} = 180^\circ$

🧠 **Sense check:** Three lone pairs ring the equator evenly. The two fluorines line up through the middle, 180° apart.



Your turn: chlorine trifluoride

ClF_3 : Cl bonded to 3 F, with 2 lone pairs on Cl
 central atom: chlorine · 3 bonded F · 2 lone pairs

step	count	result
1 · domains	3 bonded + 2 lone pairs	_____ domains
2 · geometry	from 5 domains	_____
3 · seat the lone pairs	2 corners, both _____	1 equatorial F remains
4 · shape	2 axial F + 1 equatorial F	_____

$\text{ClF}_3 \rightarrow \text{T-shaped}$

$3 + 2 = 5$ domains · trigonal bipyramidal geometry · both lone pairs equatorial · angles $\approx 90^\circ$



⚠ Where this goes wrong

Seating a lone pair axial. SF_4 's lone pair in an axial corner would leave three fluorines in a flat triangle with one on top: a trigonal pyramid. The wide pair takes an equatorial corner, and the shape is a seesaw.

Naming the electron-pair geometry as the shape. SF_4 's five domains sit trigonal bipyramidal, but one corner holds a lone pair. The shape names only the atoms: seesaw.

Forcing an octet. Treating XeF_2 like H_2O gives $2 + 2 = 4$ domains and a bent molecule. Xenon holds 2 bonds + 3 lone pairs = 5 domains, all three lone pairs equatorial: linear.

Placing octahedral lone pairs side by side. XeF_4 's two wide pairs at 90° to each other would crowd. They sit 180° apart, and the four fluorines flatten into a square plane.



Practice 1

BrF_3 : Br bonded to 3 F, with 2 lone pairs on Br
central atom: bromine · 3 bonded F · 2 lone pairs

Bromine trifluoride has three bonded fluorines and two lone pairs on bromine. What is its molecular shape?

- A. Trigonal bipyramidal: the five electron domains spread to the five corners
- B. T-shaped: both lone pairs take equatorial corners, leaving the three fluorines in a T
- C. Trigonal planar: the three fluorines spread evenly at 120°
- D. Trigonal pyramidal: three bonds and one lone pair make four domains

✓ Practice 1 answer: B

$\text{BrF}_3 \rightarrow \text{T-shaped} \rightarrow \text{answer B}$

3 bonded F + 2 lone pairs = 5 domains · trigonal bipyramidal geometry · both lone pairs equatorial

 **Sense check:** Five domains, two of them wide lone pairs in equatorial corners: the three fluorines trace a T, with angles near 90° .





Worked example 3: xenon tetrafluoride

XeF_4 : Xe bonded to 4 F, with 2 lone pairs on Xe
central atom: xenon · 4 bonded F · 2 lone pairs · every Xe-F bond polar

Xenon holds four bonded fluorines and two lone pairs. Find the molecular shape, then decide whether the molecule is polar.

✓ Worked example 3: the shape

XeF_4 : Xe bonded to 4 F, with 2 lone pairs on Xe
central atom: xenon · 4 bonded F · 2 lone pairs

Step 1 · Count the domains

Step 2 · Name the geometry

Step 3 · Seat the lone pairs

Step 4 · Name the shape

$\text{XeF}_4 \rightarrow$ square planar

6 domains · 2 lone pairs opposite each other · 4 F in one plane at 90°

🧠 **Sense check:** Two opposite corners empty out, and the four fluorines flatten into a square around the xenon.



✓ Worked example 3: the polarity

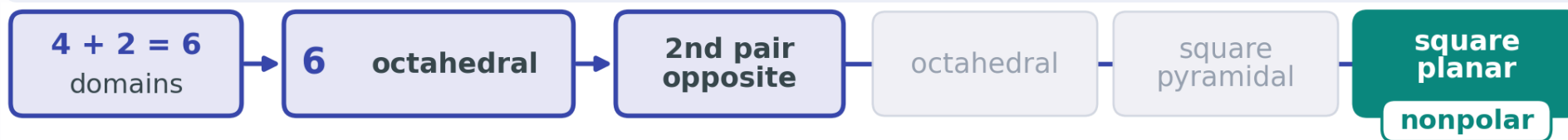
XeF₄ square planar: 4 polar Xe-F bonds
 4 equal arrows toward the corners of a square · 2 lone pairs at 180°

Polarity · Add the bond dipoles

Each Xe-F arrow has a partner pointing exactly the other way, so the four cancel in pairs. The two lone pairs sit opposite each other and balance as well.

XeF₄ → nonpolar
 4 equal arrows in a square cancel · 2 lone pairs at 180° cancel

 **Sense check:** Every arrow and both lone pairs have an exact opposite. Square planar with identical outer atoms is nonpolar.



Take-home: square planar cancels, seesaw does not

XeF₄ square planar: every arrow has a partner straight across → nonpolar
the 2 lone pairs at 180° balance as well

SF₄ seesaw: the 2 equatorial arrows share a side → polar
no arrow comes from the lone pair's corner to balance them

Symmetry decides. In a square plane every dipole has an exact opposite. A seesaw leaves two equatorial arrows unanswered, so a net dipole remains.



Practice 2

BrF_5 : Br bonded to 5 F, with 1 lone pair on Br
square pyramidal · every Br-F bond polar · all outer atoms identical

Bromine pentafluoride is square pyramidal, with five identical polar bonds. Polar or nonpolar, and why?

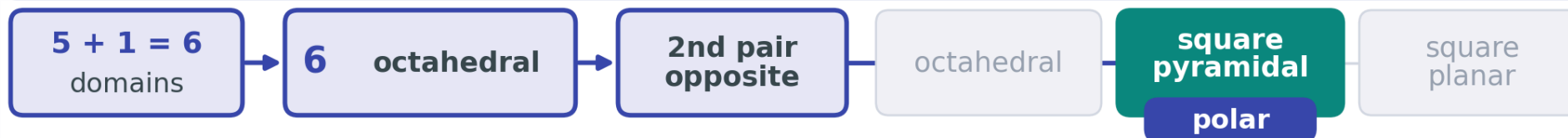
- A. Polar: the four base arrows cancel in pairs, but the fifth arrow faces the lone pair's corner and has no partner
- B. Nonpolar: five identical polar bonds always cancel, whatever the arrangement
- C. Nonpolar: the Br-F bonds are not polar in the first place
- D. Polar: any molecule that contains polar bonds is polar

✓ Practice 2 answer: A

$\text{BrF}_5 \rightarrow \text{polar} \rightarrow \text{answer A}$

4 base arrows cancel in pairs · the fifth arrow faces the lone pair · net arrow along the axis

🧠 **Sense check:** One arrow without a partner is all it takes. Square pyramidal keeps a net dipole, like seesaw and T-shaped.





Practice 3



a cation · iodine bonded to 4 F

What is the molecular shape of the IF_4^+ ion?

- A. Tetrahedral
- B. Trigonal bipyramidal
- C. Square planar
- D. Trigonal pyramidal
- E. Seesaw

✓ Practice 3 answer: E

IF_4^+ : $7 + 4(7) - 1 = 34$ · $34 - 4(2) - 4(6) = 2$ left on I → 1 lone pair
 the 1+ charge removes one electron · 4 bonded F + 1 lone pair = 5 domains · lone pair equatorial ·
 seesaw → answer E

🧠 **Sense check:** Same count as SF_4 : 34 electrons, 4 bonds, 1 lone pair. The lone pair takes an equatorial corner, and the four fluorines form a seesaw.





Check yourself

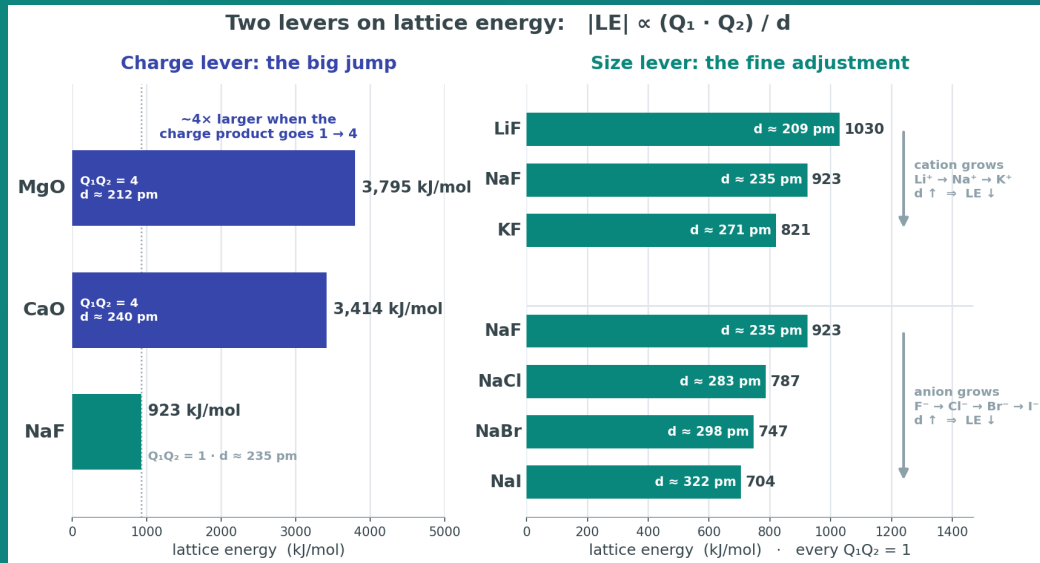
1. PCl_5 has five bonded chlorines and no lone pairs on phosphorus. Give the electron-pair geometry, the molecular shape, and both bond angles.
2. XeF_2 holds two bonded fluorines and three lone pairs. Where do the lone pairs sit, and is the molecule polar or nonpolar?

Shapes and dipoles settle how molecules meet. The closing question is strength: how firmly the bonds themselves hold. That story starts with the ionic lattice, where the grip of a whole crystal can be measured, and ends at the price of breaking one covalent bond.

8 • Lattice Energy Trends

Rank or compare ionic compounds by lattice energy using Coulomb's law (charge product first, then ion size) and connect a larger lattice energy to a higher melting point and greater hardness.

Why some salts refuse to melt



Salt melts at 801 °C; magnesium oxide holds its shape past 2800 °C. Same ionic bond, but two levers make the lattice energy wildly different.

Ionic solids: opposite charges attract

A metal gives electrons to a nonmetal, and the ions attract. The periodic table sets both things that control that pull: each ion's charge, from its group, and its size, which grows down a group.

**group 1: Na^+ · group 2: Mg^{2+} · group 16: O^{2-} · group 17: Cl^- · MgCl_2 is Mg^{2+} with Cl^-
a subscript counts ions, not charge · ions grow down a group: $\text{Li}^+ < \text{Na}^+ < \text{K}^+$ · $\text{F}^- < \text{Cl}^- < \text{Br}^- < \text{I}^-$**

Lattice energy: ionic bond strength

Bond strength is measured twice: lattice energy for ionic solids, bond energies for covalent bonds. Lattice energy is the energy that holds one mole of an ionic solid together, reported as a positive magnitude.



bigger magnitude = stronger bonding · running it in reverse (ions → solid) releases the same energy

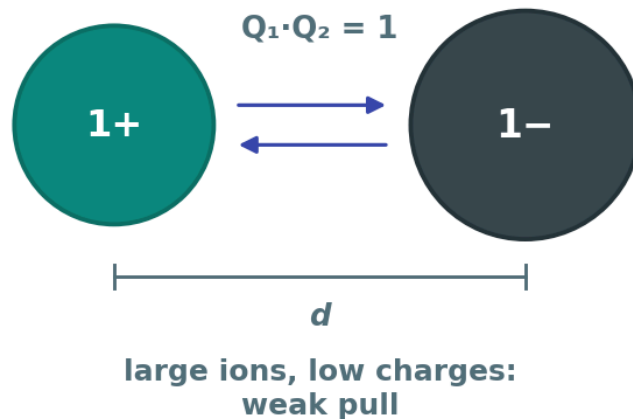
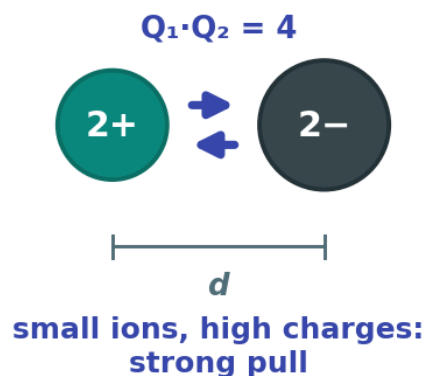
Coulomb's law sets the pull

The attraction is electrostatic, so lattice energy follows Coulomb's law: just two knobs.

$$|\text{lattice energy}| \propto (Q_1 \cdot Q_2) / d$$

Q_1, Q_2 = the ion charges · d = distance between ion centers $\approx$ sum of the two ionic radii

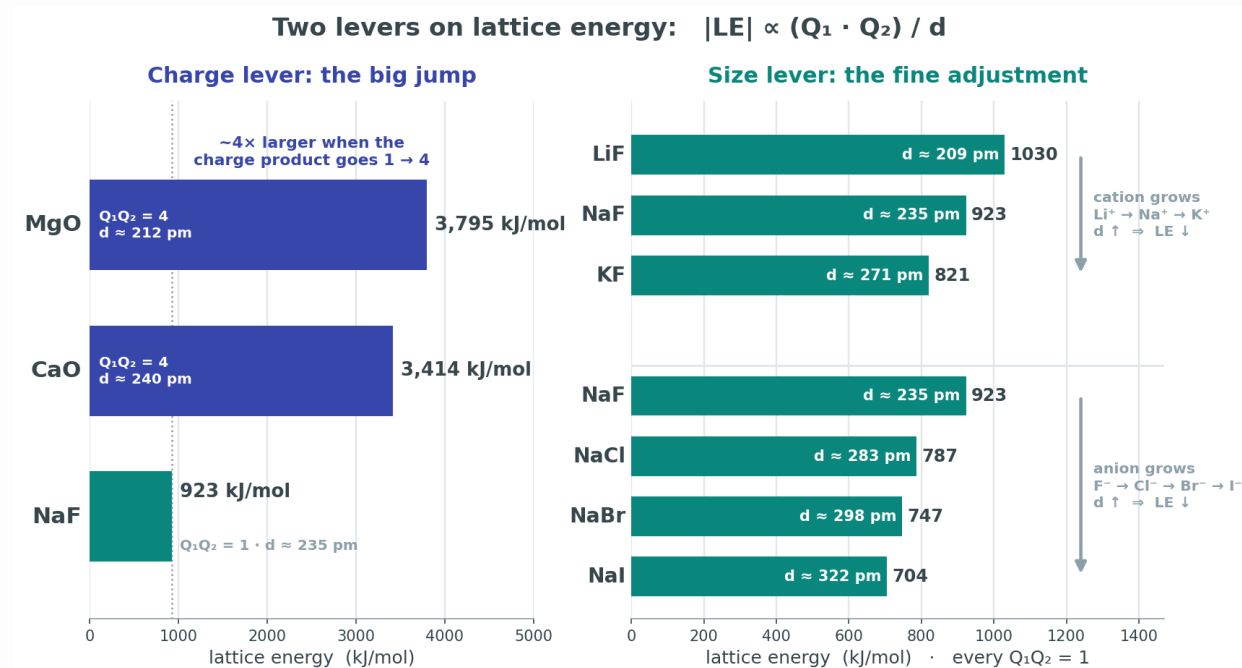
Two ion pairs, two very different pulls



The **charge product** sits on top; the **separation d** sits on the bottom. Every trend is just this equation read carefully.

Lever 1: charge is the big jump

Charge is a **product** in the numerator, so it swings hardest. Go from 1+/1- to 2+/2- and the lattice energy roughly quadruples, even at similar sizes.



MgO ($Q_1 Q_2 = 4$): 3795 $\gg$ NaF ($Q_1 Q_2 = 1$): 923 kJ/mol
3795 / 923 $\approx$ 4.1: quadruple the charge product, quadruple the pull ($d \approx 212$ vs 235 pm)

Lever 2: size is the fine adjustment

Hold the charges equal and only **d** changes. Smaller ions sit closer (shorter d, larger lattice energy) while bigger ions weaken it.

NaF 923 > NaCl 787 > NaBr 747 > NaI 704 (anion grows, d ↑)

LiF 1030 > NaF 923 > KF 821: cation grows down a group, d rises · a few hundred kJ/mol, never the fourfold charge jump

What a big lattice energy buys

A stronger lattice resists being pulled apart: by heat, by scratching, by a solvent.

higher lattice energy → higher melting point · harder · less soluble

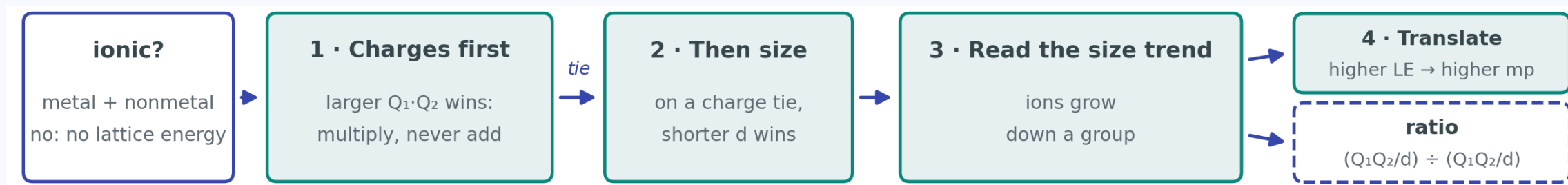
MgO melts near 2850 °C and serves as a furnace lining; NaCl melts at 801 °C

So a lattice-energy ranking is also, roughly, a melting-point and hardness ranking.



The method: compare two

1. **Charges first.** The higher $Q_1 \cdot Q_2$ product usually wins.
2. **Then size.** On a tie, smaller ions (shorter d) win.
3. **Read the size trend:** ions grow down a group.
4. **Translate:** higher lattice energy $\rightarrow$ higher melting point and hardness.





Guided example: which melts higher



wanted: the higher melting point

Both are ionic solids. Predict which one melts at the higher temperature. Run all four steps.

✓ Guided example: solution

MgO: Mg^{2+} with O^{2-} , $Q_1Q_2 = 2 \times 2 = 4$ · CaS: Ca^{2+} with S^{2-} , $Q_1Q_2 = 2 \times 2 = 4$
 wanted: the higher melting point · the charge products tie

Step 1 · Charges first

Step 2 · Then size

Step 3 · Read the size trend

MgO: $d \approx 72 + 140 = 212 \text{ pm}$ < CaS: $d \approx 100 + 184 = 284 \text{ pm}$
 shorter d , larger lattice energy: $\text{MgO} > \text{CaS}$

Step 4 · Translate

The larger lattice energy holds the solid harder, so **MgO melts higher**.

🧠 **Sense check:** Measured: MgO 2852 °C, CaS 2525 °C, a gap of 327 °C. Equal charges, so size moves it by hundreds of degrees, not thousands. ✓





Worked example 1

Rank by lattice energy, largest first: NaBr, MgO, NaF
then estimate how many times larger MgO's lattice energy is than NaF's

MgO is built from Mg^{2+} and O^{2-} ; NaF and NaBr are each $1+/1-$. Apply the charge lever first, then break the tie with size.

✓ Worked example 1: solution

Step 1 • Charges first

Step 2 • Then size

MgO (3795) > NaF (923) > NaBr (747)

charge sets the leader, size orders the rest

Estimate the jump

$$|LE|(MgO) / |LE|(NaF) \approx (4/212) / (1/235) \approx 4.4$$

measured $3795 / 923 \approx 4.1$: quadrupling the charge product roughly quadruples the lattice energy

🧠 **Sense check:** Charge moved the answer by a factor of ~4; size only reorders the two 1+/1– salts by a few percent.





Worked example 2

LiF has the smallest ions of any alkali halide.

Does that give it a larger lattice energy than CaO?

Small ions are tempting: a short d means a strong pull. But check the other lever before deciding.

✓ Worked example 2: solution

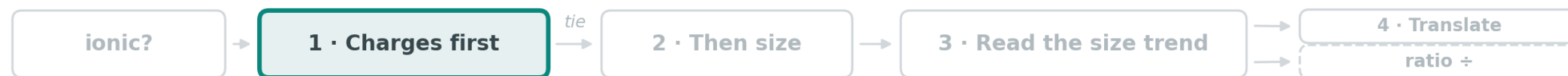
Step 1 · Charges first

Size does not apply

LiF's ions are smaller ($d \approx 209$ vs 240 pm), but size only breaks a charge tie. Here 4 beats 1 decisively, so the route stops at Step 1.

CaO (3414 kJ/mol) > LiF (1030 kJ/mol) $\approx 3.3\times$ larger
 $3414 / 1030 \approx 3.3$: the fourfold charge edge outweighs LiF's small size advantage

 **Sense check:** Smallest ions $\neq$ largest lattice energy. Size is the fine adjustment; charge is the main lever.



Your turn: rank three salts

Rank **KCl**, **CaO**, **KF** by lattice energy, largest first. Fill each rank, using charge first and then size.

compound	charge product	rank (1 = largest)
CaO	_____	_____
KF	_____	_____
KCl	_____	_____



CaO wins on charge · KF beats KCl on size, since F^- is smaller than Cl^- (shorter d)



⚠ Where this goes wrong

Letting size beat charge. LiF has the smallest ions of the alkali halides, yet CaO ($2+/2-$) still crushes it: charge is the stronger lever. Check $Q_1 \cdot Q_2$ first.

Bigger ion, bigger energy. Backwards. A larger ion means a larger d , and d is in the denominator, so the lattice energy gets smaller, not larger.

Adding charges instead of multiplying. It is the product $Q_1 \cdot Q_2$, not the sum. Going $1+/1- \rightarrow 2+/2-$ multiplies the product by 4 ($1 \rightarrow 4$), not by 2.

Reading the sign, not the size. Whether you call it $+923$ or -923 , compare magnitudes: the bigger number is the stronger lattice.

Practice 1

Which of these ionic solids has the **highest** lattice energy?

- A. BaO
- B. Na₂O
- C. LiCl
- D. KI

✓ Practice 1 answer: A

BaO: $2 \times 2 = 4$ · **Na₂O:** $1 \times 2 = 2$ · **LiCl:** $1 \times 1 = 1$ · **KI:** $1 \times 1 = 1$
 $Q_1 \cdot Q_2$ for each · only BaO has both ions doubly charged → answer A

B let one 2– ion carry it: Na₂O is Na⁺ with O^{2–}, so $Q_1 Q_2 = 1 \times 2 = 2$, half of BaO's 4. C let size beat charge: LiCl's ions sit closer than BaO's ($d \approx 257$ vs 275 pm), but its product is 1. D ran size backward: KI has the largest ions here, and a longer d lowers the lattice energy.

🧠 **Sense check:** Ba²⁺ is a large ion, yet BaO still wins. Charge first, size second. ✓





Practice 2

KBr · LiBr

wanted: the greater lattice energy

Which has the greater lattice energy, KBr or LiBr?

- A. KBr: its larger cation attracts Br^- more strongly
- B. LiBr: Li^+ carries a higher charge than K^+
- C. LiBr: equal charges, and the smaller Li^+ shortens d
- D. Equal: both are $1+/1-$ bromides
- E. KBr: its larger formula mass holds it together

✓ Practice 2 answer: C

LiBr: $Q_1Q_2 = 1$, $d \approx 76 + 196 = 272$ pm · KBr: $Q_1Q_2 = 1$, $d \approx 138 + 196 = 334$ pm
charges tie · Li sits above K in group 1, so Li^+ is smaller · shorter d wins → answer C

A ran size backward: a larger cation means a longer d , and d sits in the denominator. B misread the charge: every group 1 ion is $1+$, whatever its row. D stopped at the tie: equal charges hand the decision to size. E credited mass: formula mass is not in Q_1Q_2/d .

🧠 **Sense check:** A tie on charge is not a tie on lattice energy. LiBr's d is 62 pm shorter, so LiBr wins. ✓





Practice 3

CaS: $d = 284 \text{ pm}$ · KCl: $d = 319 \text{ pm}$
wanted: $|LE|(\text{CaS}) \div |LE|(\text{KCl})$

The lattice energy of CaS is about how many times that of KCl?

- A. 2.25
- B. 3.56
- C. 4.00
- D. 4.49
- E. 1.12

✓ Practice 3 answer: D

$$(2 \times 2 / 284) \div (1 \times 1 / 319) = 4 \times 319/284 = 4.49 \rightarrow \text{answer D}$$

CaS is $\text{Ca}^{2+} \text{S}^{2-}$, so $Q_1 Q_2 = 4$ · KCl is $\text{K}^+ \text{Cl}^-$, so $Q_1 Q_2 = 1$ · charge product on top, d on the bottom

A added the charges: $(2 + 2)/(1 + 1) \times 319/284 = 2.25$. B flipped the distance ratio: $4 \times 284/319 = 3.56$, which rewards the longer d. C stopped at the charge product: $4/1 = 4.00$, with size never applied. E skipped the charge lever: $319/284 = 1.12$, the distance ratio alone.

🧠 **Sense check:** Charge sets a factor of 4; CaS's shorter d pushes it a little above 4, since d sits in the denominator.



Check yourself

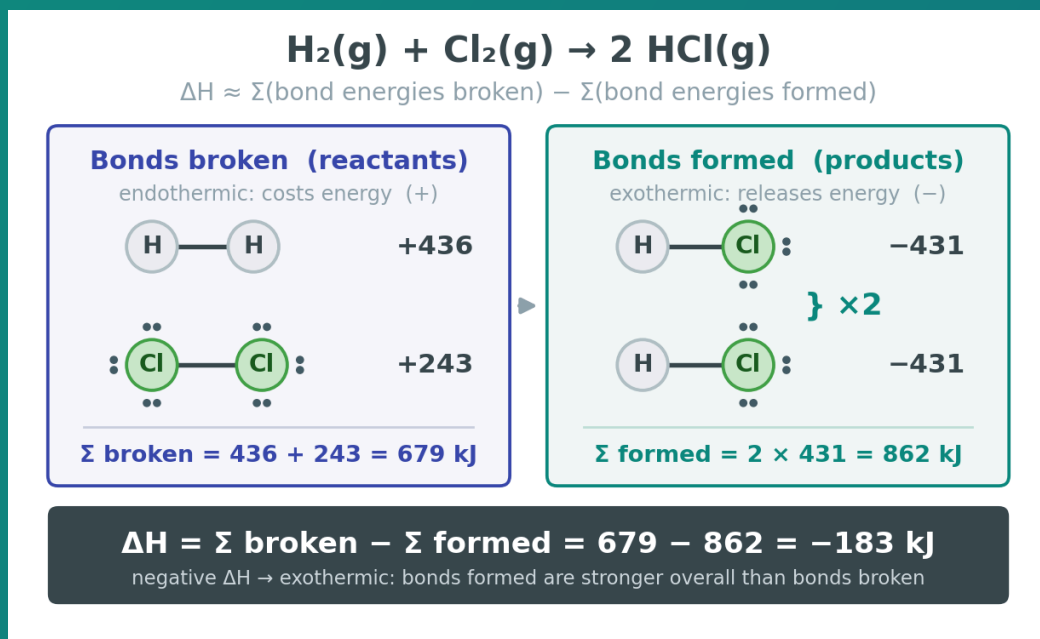
1. Which has the larger lattice energy, KF or KI? Which lever did you use, and which melts higher?
2. A salt's ions both change from $1+/1-$ to $3+/3-$ at the same spacing. By what factor does the lattice energy change?

Charge and size explain why ionic solids differ so widely in melting point and hardness. Lattice energy measures ionic bond strength; bond energies measure the covalent kind, and estimating a reaction's ΔH with them is the next step.

9 · Bond Energies & Reaction Enthalpy

Estimate the enthalpy change of any balanced gas-phase reaction by counting every bond from Lewis structures and applying $\Delta H \approx \Sigma(\text{bond energies broken}) - \Sigma(\text{bond energies formed})$.

Bonds broken, bonds formed



Every reaction tears old bonds apart and stitches new ones together. Breaking costs energy; forming pays it back. Compare the two totals and you know the heat before running the reaction.

Hess's law, run through free atoms

(1) reactants $\rightarrow$ free gas atoms $\cdot \Delta H_1 > 0$: every reactant bond breaks

(2) free gas atoms $\rightarrow$ products $\cdot \Delta H_2 < 0$: every product bond forms

Hess's law: $\Delta H = \Delta H_1 + \Delta H_2$ $\cdot$ any path gives the same ΔH $\cdot$ Lewis structures list the bonds on each leg

Hess's law lets any path stand in for the real one. Take the path through free atoms: pull every reactant apart, then build every product. The legs carry opposite signs; ΔH is their balance.

Bond energy: the price of one bond

D is the energy to break **one mole** of a bond in the **gas phase**. Bonded atoms sit lower in energy than separated ones, so pulling them apart is always a climb.

breaking a bond → ENDOTHERMIC (+, costs energy)

forming a bond → EXOTHERMIC (–, releases energy) · same size, opposite sign

breaking H-H costs +436 kJ/mol · forming H-H releases –436 kJ/mol

memory hook: breaking costs, forming pays · one number used both ways · D itself is always positive

The master formula

A reaction breaks every reactant bond, then forms every product bond. Sum the D values on each side.

$$\Delta H \approx \sum D(\text{bonds broken}) - \sum D(\text{bonds formed})$$

reactant bonds cost (+) · product bonds pay (–) · negative $\Delta H \rightarrow$ exothermic



a double or triple bond is ONE bond with its own larger D · never a stack of singles

Count from the Lewis structures: a bond you did not draw is a bond you will not tally.

The method

1. **Draw every Lewis structure.** Balanced equation, gas phase.
2. **Tally bonds broken.** Every reactant bond $\times$ its coefficient; sum the D values.
3. **Tally bonds formed.** Every product bond $\times$ its coefficient.
4. **Subtract.** $\Delta H = \Sigma \text{ broken} - \Sigma \text{ formed}$.



 Worked example 1: $\text{H}_2 + \text{Cl}_2 \rightarrow 2 \text{HCl}$ 

$$D: \text{H-H} = 436 \cdot \text{Cl-Cl} = 243 \cdot \text{H-Cl} = 431 \text{ (kJ/mol)}$$

Reactants hold one H-H bond and one Cl-Cl bond. The products are **2** HCl: two H-Cl bonds. Break the reactant bonds, form the product bonds, and compare.

✓ Worked example 1: solution



D: H-H = 436 · Cl-Cl = 243 · H-Cl = 431 (kJ/mol)

Step 1 · Draw every Lewis structure

Step 2 · Tally bonds broken

Step 3 · Tally bonds formed

Step 4 · Subtract

$$\Delta H = 679 - 862 = -183 \text{ kJ}$$

negative → exothermic · the two H-Cl bonds release more than the H-H and Cl-Cl cost

🧠 **Sense check:** Breaking costs 679 kJ; forming pays 862 kJ back. The payback wins by 183 kJ, released as heat.



 Worked example 2: $2 \text{H}_2 + \text{O}_2 \rightarrow 2 \text{H}_2\text{O}$ 

$$D: \text{H-H} = 436 \cdot \text{O=O} = 498 \cdot \text{O-H} = 467 \text{ (kJ/mol)}$$

Two traps live here: the coefficient **2** on both H_2 and H_2O , and the **O=O double bond** (one bond, $D = 498$). Each water molecule holds **two** O-H bonds. Estimate ΔH .

✓ Worked example 2: solution



D: H-H = 436 · O=O = 498 · O-H = 467 (kJ/mol)

Step 1 · Draw every Lewis structure

Step 2 · Tally bonds broken

Step 3 · Tally bonds formed

Step 4 · Subtract

$$\Delta H = 1370 - 1868 = -498 \text{ kJ}$$

exothermic: this is why hydrogen burns · measured with water as a gas $\approx -484 \text{ kJ}$, close

 **Sense check:** Both traps defused: the O=O kept its own single D value, and the coefficients turned two waters into four O-H bonds.



Take-home: coefficients multiply the bond count

$2 \text{ H}_2\text{O} \rightarrow 2 \times 2 = 4 \text{ O-H bonds formed}$

$\Sigma \text{ formed} = 4 \times 467 = 1868 \text{ kJ} \checkmark \cdot \Delta H = 1370 - 1868 = -498 \text{ kJ}$

counting one molecule's bonds: $\Sigma \text{ formed} = 2 \times 467 = 934 \text{ kJ} \times$

$\Delta H = 1370 - 934 = +436 \text{ kJ} \cdot$ the estimate lands on the wrong side of zero

A coefficient multiplies every bond in its molecule. Miss one and a reaction that burns comes out endothermic on paper.

Why the answer is an estimate

The D values are **averages**: the true C-H in methane differs a little from the C-H in ethanol. And they are gas-phase numbers, while real reactions may involve liquids and solids.

bond-energy $\Delta H \approx$ measured ΔH

averages + gas-phase assumption $\rightarrow$ an estimate, usually within a few percent · plenty to call exothermic vs endothermic

Your turn: $\text{H}_2 + \text{Br}_2 \rightarrow 2 \text{HBr}$



D: H-H = 436 · Br-Br = 193 · H-Br = 366 (kJ/mol)

step	tally	result
1 · draw every Lewis structure	all single bonds	H-H · Br-Br · 2 × H-Br
2 · tally bonds broken	436 + 193	_____ kJ
3 · tally bonds formed	2 × 366	_____ kJ
4 · subtract	Σ broken – Σ formed	ΔH = _____ kJ

Fill the three results, then call it: exothermic or endothermic?

$$\Delta H = 629 - 732 = -103 \text{ kJ} \rightarrow \text{exothermic}$$

broken: $436 + 193 = 629$ · formed: $2 \times 366 = 732$ · forming pays more than breaking costs

⚠ Where this goes wrong

Flipping the formula. It is Σ broken – Σ formed, not formed – broken. Reversing it flips the sign and turns exothermic into endothermic.

Ignoring coefficients. 2 HCl means two H-Cl bonds; 2 H₂O means four O-H bonds. Multiply each bond count by its coefficient.

Using a single-bond value for a multiple bond. O=O is 498 and N≡N is 941: one bond each, with their own D. Never split them into singles.

Skipping the Lewis structures. You cannot tally bonds you did not draw. Sketch every molecule first, then count.

 Practice 1

D: O=O = 498 · F-F = 155 · O-F = 190. (kJ/mol)

Estimate ΔH , in kJ, for this reaction.

- A. -760.
- B. +48
- C. +428
- D. -48

✓ Practice 1 answer: B

$$\Delta H = (498 + 2 \times 155) - 2 \times 2 \times 190. = 808 - 760. = +48 \text{ kJ} \rightarrow \text{answer B}$$

break O=O + 2 F-F = 498 + 310 = 808 · form 2 OF₂ × 2 O-F = 4 O-F = 4 × 190. = 760.

A (−760.) summed only the bonds formed; the 808 kJ of breaking was never charged. C (+428) forgot the coefficient 2 on OF₂: 808 − 380. = 428. D (−48) reversed the formula: 760. − 808 = −48, the right size with the wrong sign.

🧠 **Sense check:** Endothermic: the four weak O-F bonds do not repay the O=O and the two F-F. ✓



 Practice 2

D: C-H = 413 · C=C = 611 · O=O = 498 · C=O in CO₂ = 799 · O-H = 467 (kJ/mol)

Ethylene burns. Estimate ΔH , in kJ, for this reaction.

- A. -1307
- B. +1307
- C. +1225
- D. -5064
- E. +3757

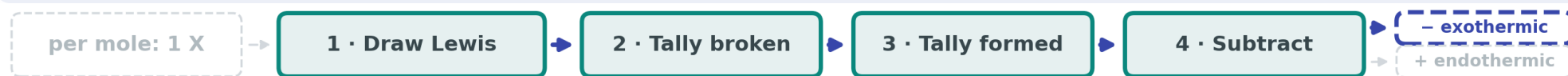
✓ Practice 2 answer: A

$$\Delta H = (1652 + 611 + 1494) - (3196 + 1868) = 3757 - 5064 = -1307 \text{ kJ} \rightarrow \text{answer A}$$

broken: 4 C-H (1652) + 1 C=C (611) + 3 O=O (3 × 498 = 1494) = 3757 · formed: 4 C=O (4 × 799 = 3196) + 4 O-H (1868) = 5064

B (+1307) reversed the formula: 5064 – 3757. C (+1225) ignored both product coefficients, one CO₂ and one H₂O: 1598 + 2 × 467 = 2532, and 3757 – 2532 = +1225. D (–5064) summed the products only. E (+3757) stopped at Σ broken and never subtracted.

 **Sense check:** Every tally carries a coefficient: 3 O=O broken, 2 CO₂ and 2 H₂O formed. Payback beats cost by 1307 kJ, which is why ethylene burns hot. ✓





Guided example: ammonia as a fuel



unbalanced · D: N-H = 391 · O=O = 498 · N≡N = 941 · O-H = 467 (kJ/mol)

Cargo ships are testing ammonia as a carbon-free fuel. Estimate ΔH , in kJ per mole of NH_3 , for its combustion.

✓ Guided example: solution



N: 1 = 1 · H: 3 = 3 · O: 3/2 = 3/2 · D: N-H = 391 · O=O = 498 · N≡N = 941 · O-H = 467 (kJ/mol)

Step 1 · Draw every Lewis structure

Step 2 · Tally bonds broken

Step 3 · Tally bonds formed

Step 4 · Subtract

$$\Delta H = 1546.5 - 1871.5 = -325 \text{ kJ per mole of NH}_3$$

exothermic · measured with water as a gas $\approx -317 \text{ kJ/mol}$, close · $4 \text{NH}_3 + 3 \text{O}_2 \rightarrow 2 \text{N}_2 + 6 \text{H}_2\text{O}$ burns 4 mol: $4 \times (-325) = -1300 \text{ kJ}$

 **Sense check:** Per mole of fuel fixes the 1 in front of NH_3 . Every other coefficient, fractions included, multiplies its bonds. ✓



 Practice 3

unbalanced · D: C-H = 413 · C-C = 347 · C=C = 611 · O=O = 498 · C=O in CO₂ = 799 · O-H = 467
(kJ/mol)

A propene leak catches fire at a refinery. What is ΔH , in kJ per mole of propene burned?

- A. -3662
- B. -3838
- C. -1919
- D. -4160.

✓ Practice 3 answer: C



$$\text{C: } 3 = 3 \cdot \text{H: } 6 = 6 \cdot \text{O: } 9 = 9 \cdot \text{broken: } 6 \text{ C-H (2478)} + 1 \text{ C-C (347)} + 1 \text{ C=C (611)} + 9/2 \text{ O=O (2241)} = 5677 \cdot \text{formed: } 6 \text{ C=O (4794)} + 6 \text{ O-H (2802)} = 7596$$

A (−3662) left O₂ at 1, so one O=O broke: 3436 + 498 = 3934, and 3934 − 7596 = −3662.
 B (−3838) is ΔH for 2 CH₂CHCH₃ + 9 O₂ → 6 CO₂ + 6 H₂O, two moles of propene. D (−4160.) never broke the O=O bonds: 3436 − 7596.

 **Sense check:** Measured, with water as a gas: about −1926 kJ/mol, close. Per mole of fuel puts a 1 in front of propene; 9/2 mol of O=O still breaks. ✓



Check yourself

1. For $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightarrow 2 \text{NH}_3(\text{g})$, which bonds break and which form, and how many of each? (You would need $D(\text{N}\equiv\text{N})$, $D(\text{H}-\text{H})$, $D(\text{N}-\text{H})$.)
2. A reaction has $\Sigma \text{ broken} = 1500 \text{ kJ}$ and $\Sigma \text{ formed} = 1350 \text{ kJ}$. Exothermic or endothermic, and what is ΔH ?

One model carried this whole unit: electrons shared or transferred, counted by Lewis structures. The count set the shapes, the shapes set the polarity, and here it priced the energy: breaking costs, forming pays. Count the electrons, and the rest follows.

Can you...?

- ☐ classify a bond as ionic, polar covalent, or nonpolar covalent from the electronegativity difference?
- ☐ write the Lewis (electron dot) symbol of an atom or a monatomic ion?
- ☐ count valence electrons and draw valid Lewis structures for molecules and polyatomic ions?
- ☐ draw all resonance structures of a molecule or ion and describe the blend they represent?
- ☐ assign formal charges and use them to choose the best Lewis structure?
- ☐ draw the octet-rule exceptions: incomplete octets, odd-electron molecules, and expanded octets?
- ☐ apply VSEPR theory to predict electron-pair geometry and molecular shape?
- ☐ combine bond dipoles with molecular shape to judge whether a molecule is polar?
- ☐ assign trigonal bipyramidal and octahedral electron geometries and their lone-pair shapes, and judge the polarity of five- and six-domain molecules?

If any box stays empty, the practice site has a drill for it. 