

Chemical Bonding

Preparation for General Chemistry · Dr. Karmach

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

- Classify a compound's bonding as ionic, covalent, or both, and a bond as nonpolar covalent, polar covalent, or ionic from the electronegativity difference
- Write Lewis symbols for atoms and ions and show the electron transfer that gives an ionic compound its formula
- Draw Lewis structures for molecules and polyatomic ions, including multiple bonds, resonance forms, and the common octet exceptions
- Use VSEPR to predict the shape and bond angle of a molecule with up to four electron domains
- Combine bond polarity with shape to decide whether a molecule is polar

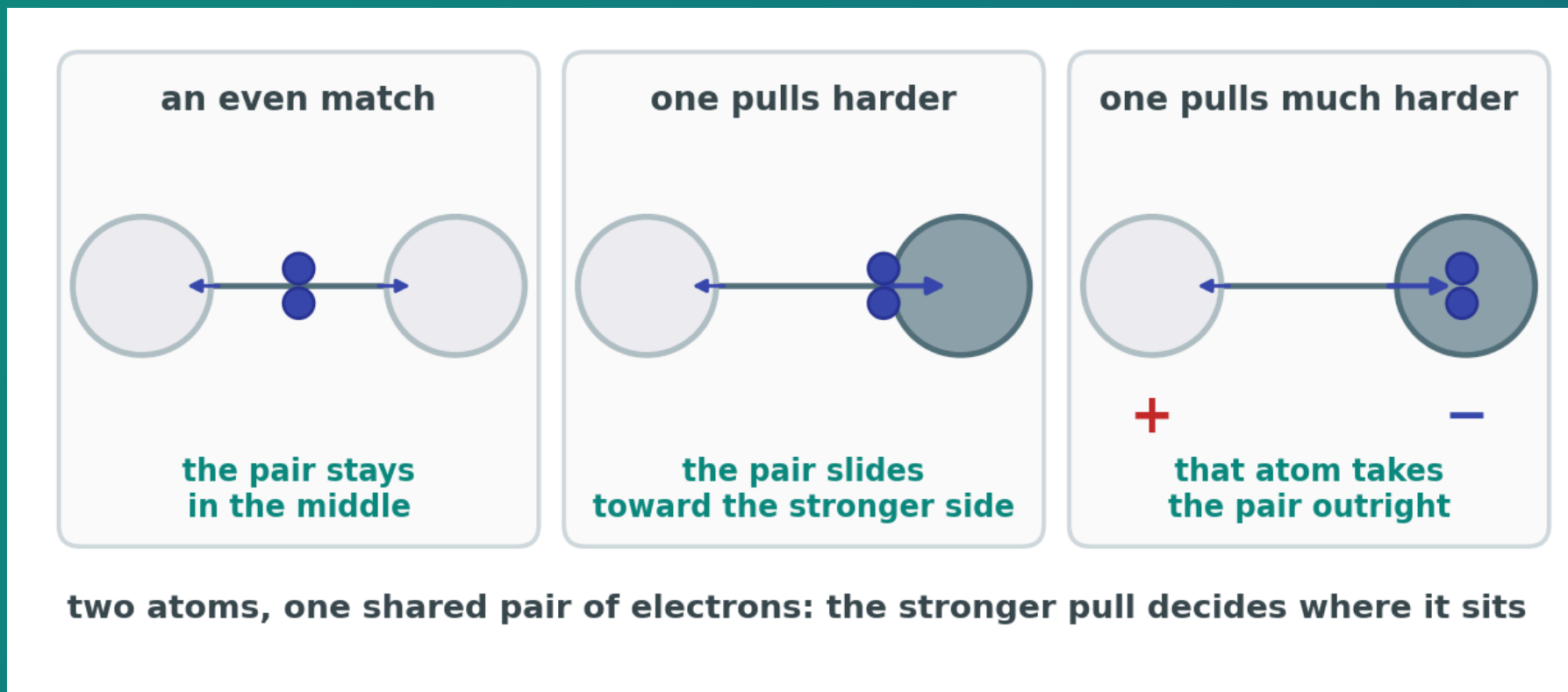
Today's route

1. **Electronegativity & Bond Type**
2. **Lewis Structures**
3. **Exceptions to the Octet Rule**
4. **VSEPR & Molecular Shape**
5. **Molecular Polarity**

1 · Electronegativity & Bond Type

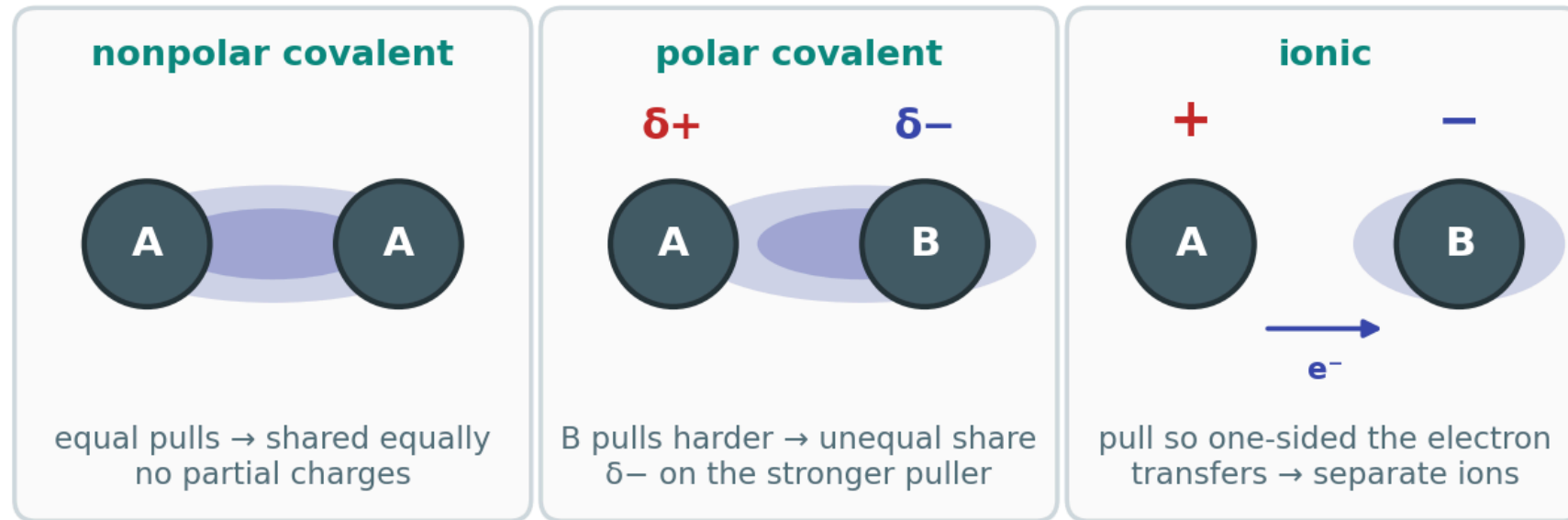
Classify any bond as nonpolar covalent, polar covalent, or ionic from the electronegativity difference, and mark which atom carries the partial negative charge, then classify a whole compound's bonding as ionic, covalent, or both.

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.

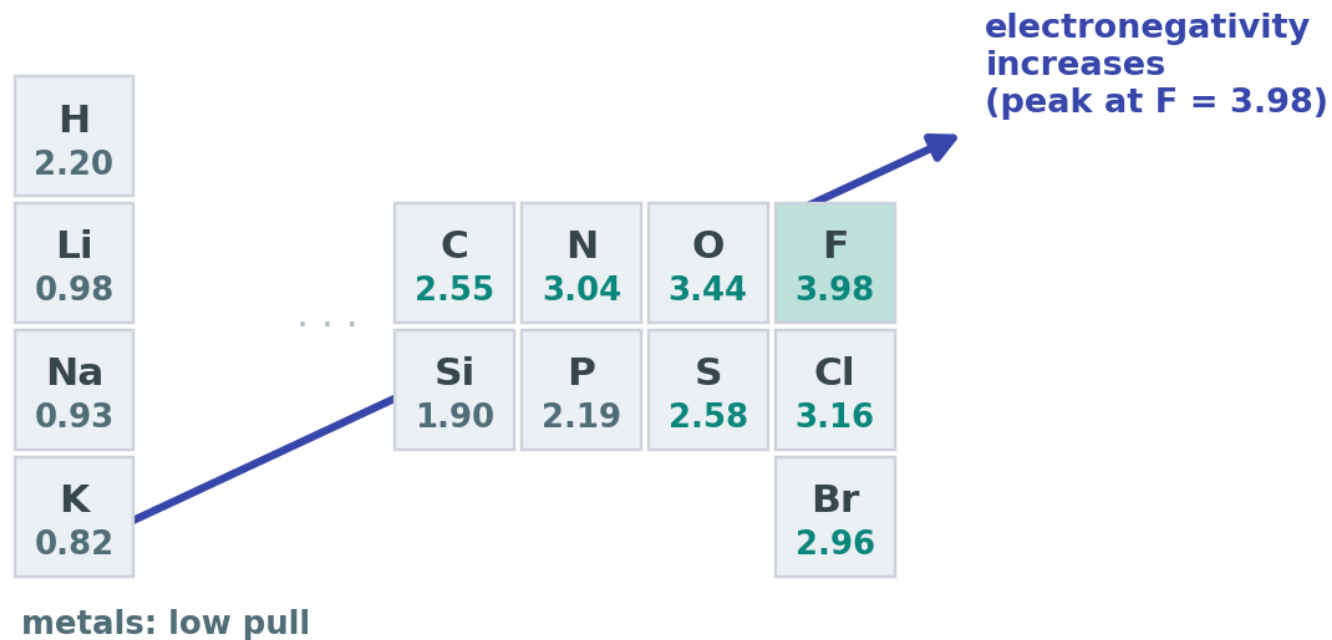
How a bond shares its electrons



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

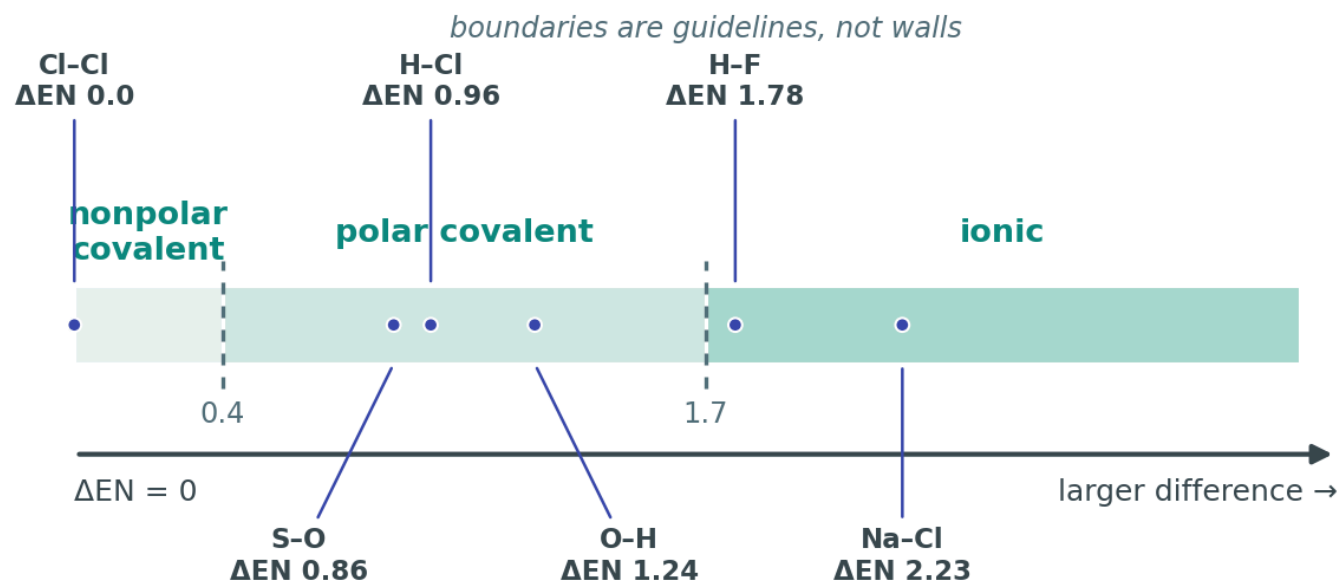
A bond is a shared pair of electrons. Each atom pulls on that pair; the strength of the pull is its electronegativity. Equal pulls share equally; unequal pulls share unequally.

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

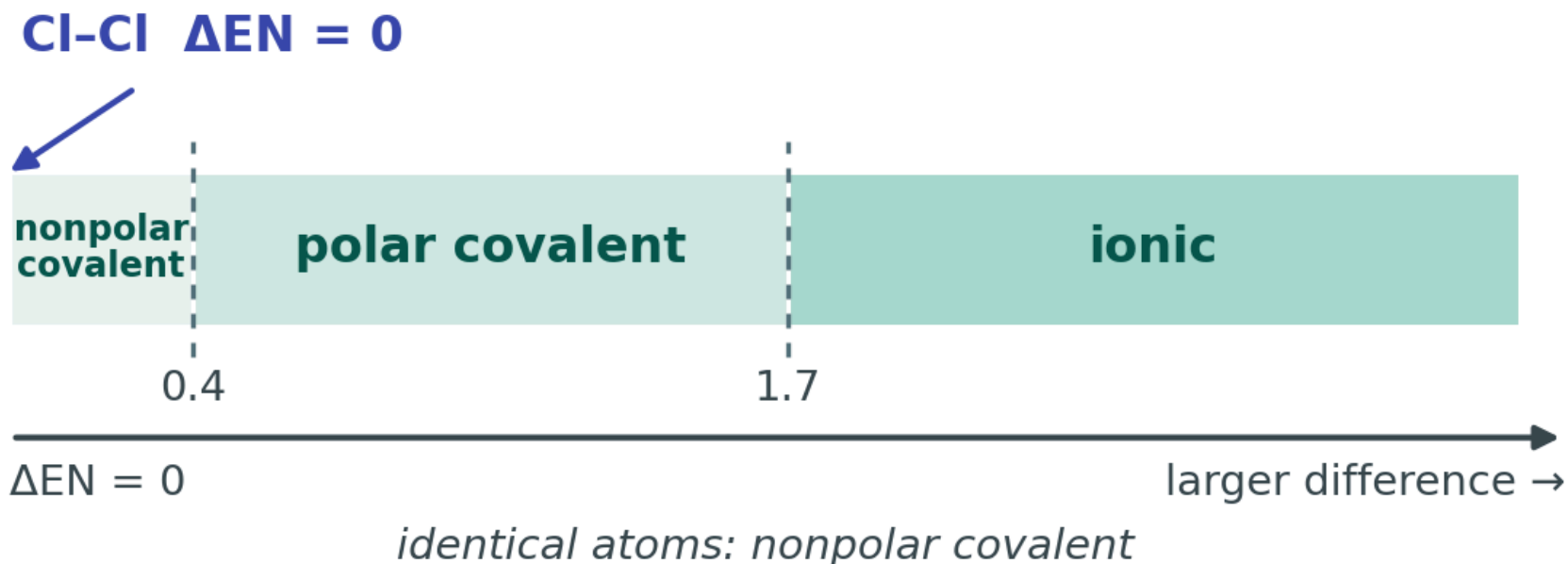
Cl-Cl → nonpolar covalent
equal electronegativities · the pair sits centered · no partial charges

 **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

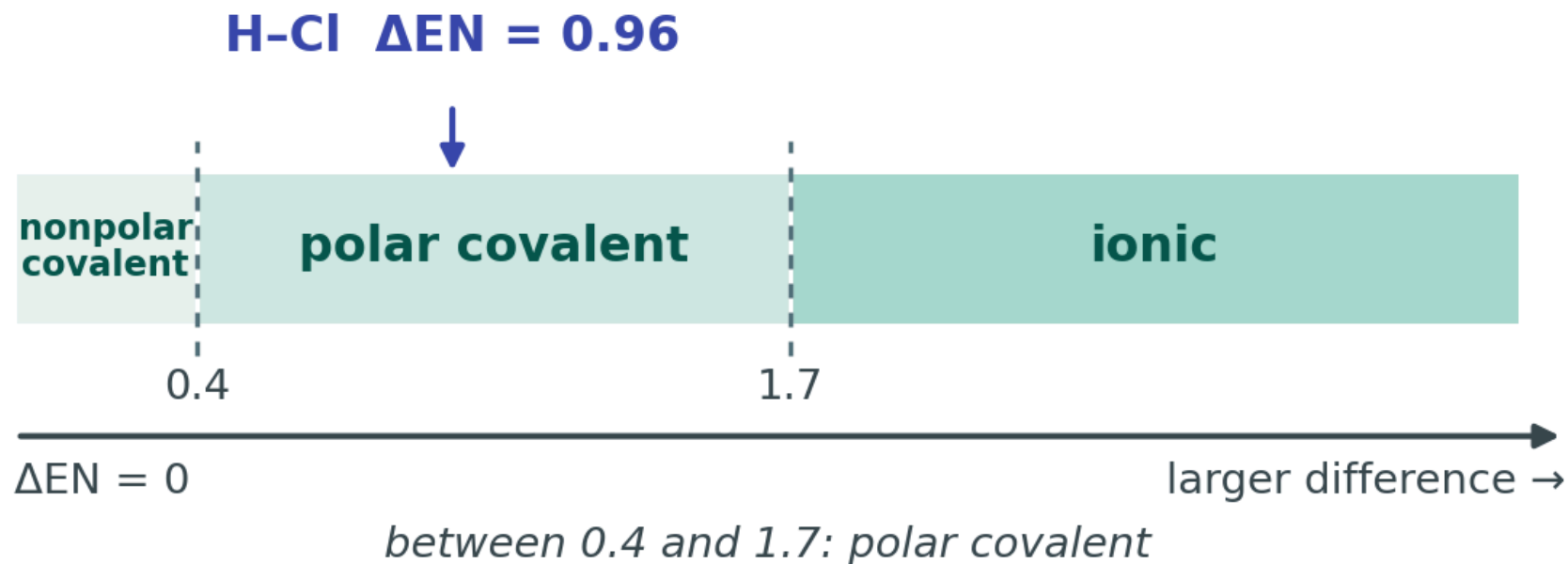
H-Cl → polar covalent, δ^- on Cl
Cl (3.16) pulls harder than H (2.20) · the shared pair shifts toward Cl · H is δ^+

 **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.

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

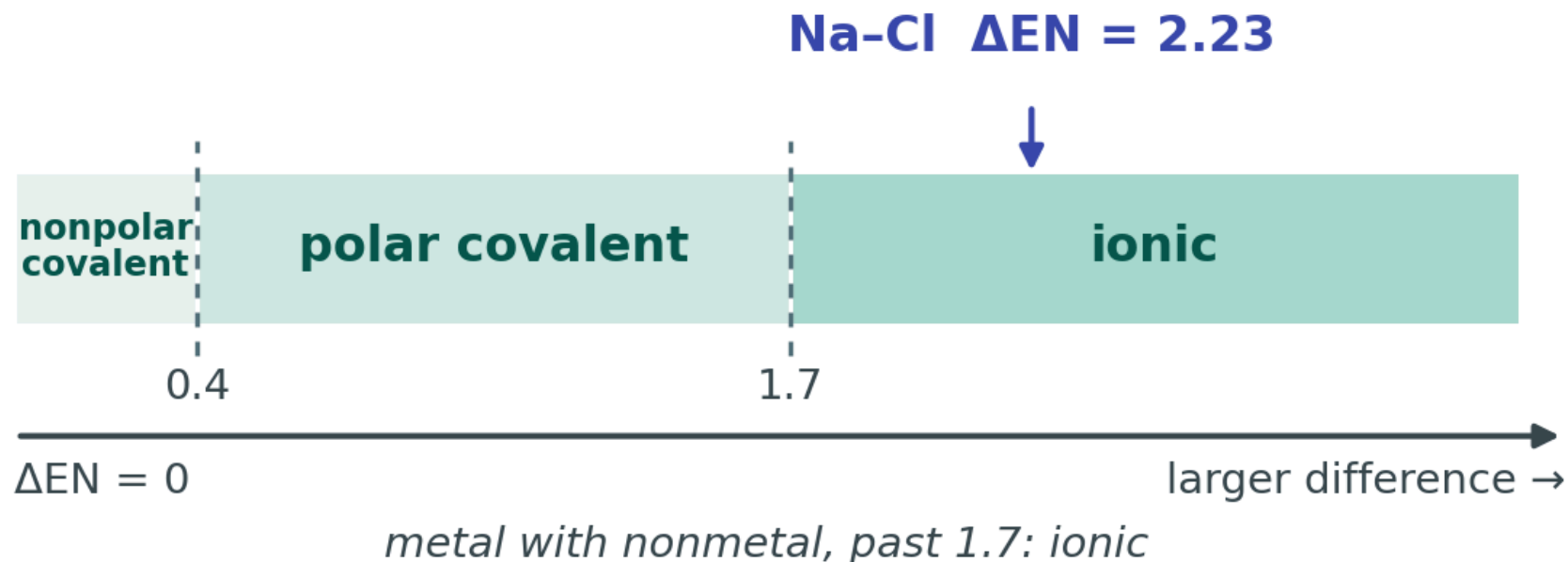
Na-Cl → ionic

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

✓ Your turn: the bond on the continuum

$$\text{Na-Cl: } \Delta\text{EN} = 3.16 - 0.93 = 2.23$$

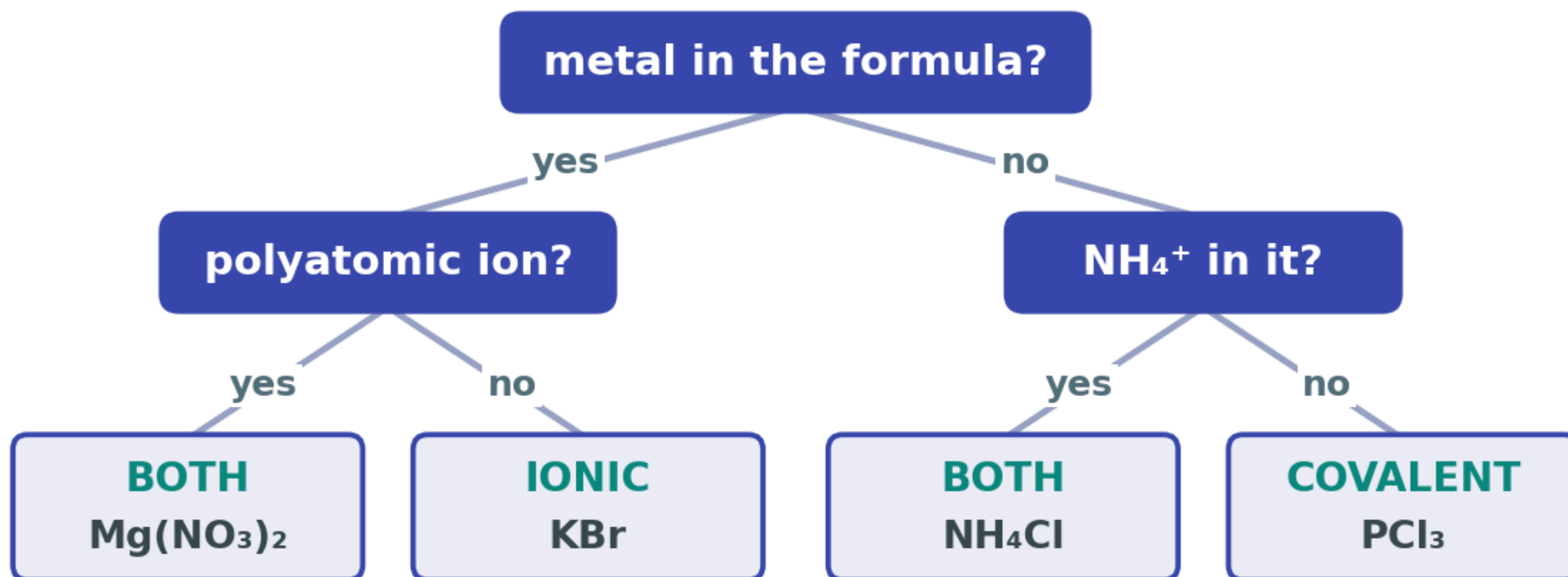
electronegativity: Na = 0.93 · Cl = 3.16 · found: ionic



 **Sense check:** 2.23 lands well past 1.7, and Na is a metal: the electron transfers.

Ionic, covalent, or both

A compound's bonding follows from its parts. Metal with nonmetal: ionic. Nonmetals only: covalent. A polyatomic ion brings both: atoms inside the ion share electrons, and ionic attraction holds the whole ion to its partner.



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: bonding in four compounds



wanted: ionic, covalent, or both, for each compound in order

What bonding does each compound have, in order?

- A. ionic · covalent · both · both
- B. ionic · covalent · ionic · ionic
- C. ionic · covalent · both · ionic
- D. ionic · covalent · covalent · covalent
- E. ionic · ionic · both · both

✓ Practice 1 answer: A



Na, Fe, Ca: metals · Fe^{2+} + SO_4^{2-} · Ca^{2+} + 2 ClO_2^- · SO_4^{2-} and ClO_2^- : polyatomic ions

$\text{NaF} \rightarrow$ ionic · $\text{SO}_2 \rightarrow$ covalent · $\text{FeSO}_4 \rightarrow$ both · $\text{Ca}(\text{ClO}_2)_2 \rightarrow$ both $\rightarrow$ **answer A**

B stopped at the metal; S-O inside SO_4^{2-} and Cl-O inside ClO_2^- are shared pairs. C missed the chlorite ion: ClO_2^- is chlorate, ClO_3^- , minus one oxygen, held together by shared Cl-O pairs. D kept only the bonds inside the ions; Fe^{2+} attracting SO_4^{2-} is ionic bonding. E read SO_2 as ions; S and O are both nonmetals, so they share: covalent only.

 **Sense check:** A metal with a polyatomic ion gives both: shared pairs inside the ion, ionic attraction between the ions.



Practice 2: ionic, covalent, or both



wanted: the compound with both ionic and covalent bonding

Which compound contains both ionic and covalent bonds?

- A. CaCl_2
- B. NH_4NO_3
- C. P_2O_5
- D. CH_3OH

✓ Practice 2 answer: B



→ N-H and N-O shared inside each ion · NH_4^+ attracts NO_3^- → both → **answer B**

A counted the two chlorines as bonded to each other; CaCl_2 is Ca^{2+} with two separate Cl^- ions, ionic only. C read the subscripts as a formula unit of ions; P and O are both nonmetals, so P_2O_5 shares: covalent only. D read the OH as hydroxide; CH_3OH has no ions at all, every atom a nonmetal sharing: covalent only.

 **Sense check:** No metal does not mean no ions. NH_4^+ is a cation built of nonmetals, so any compound of it has covalent bonds inside the ion and ionic attraction between ions.

Practice 3: 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 3 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

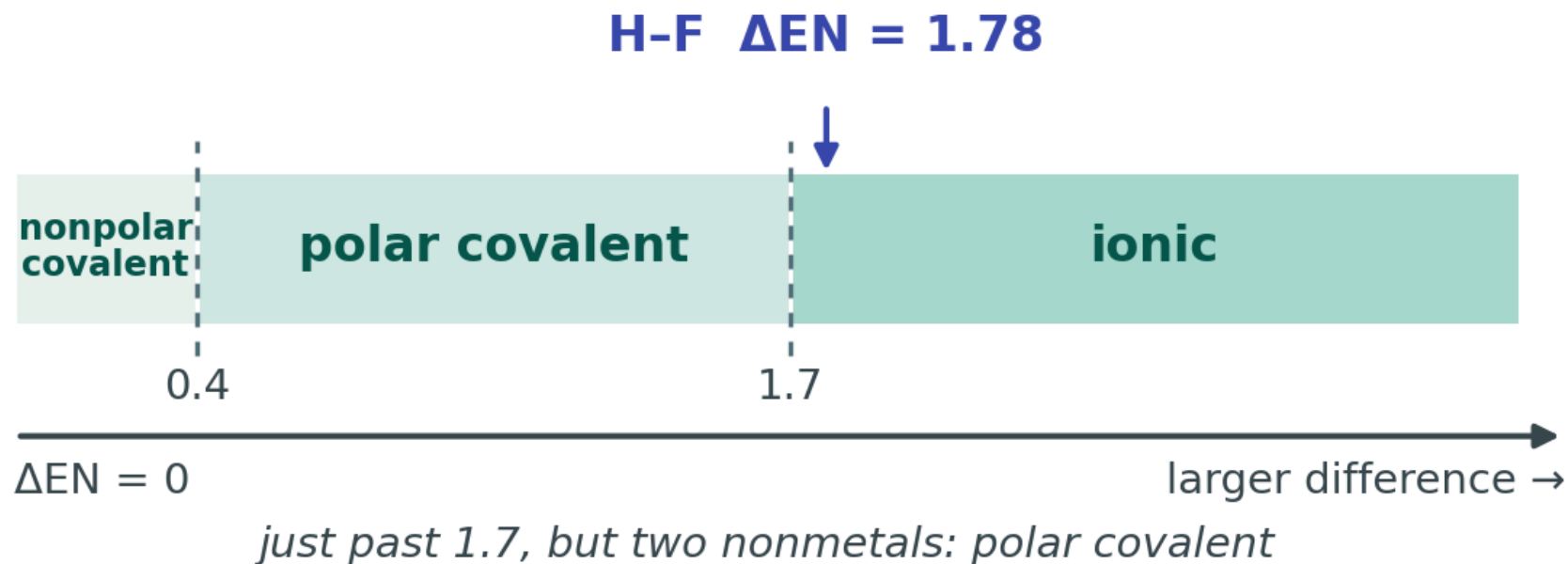
H-F → polar covalent, δ^- on F
two nonmetals · the pair stays shared but pulled far toward F · H is δ^+

 **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 4: 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 4 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^- .

Practice 5: four bonds to fluorine



electronegativity: Li = 0.98 · H = 2.20 · N = 3.04 · F = 3.98 · wanted: each bond's type, in order

What type is each bond, in order?

- A. ionic · ionic · ionic · nonpolar covalent
- B. ionic · polar covalent · nonpolar covalent · nonpolar covalent
- C. ionic · polar covalent · polar covalent · polar covalent
- D. ionic · polar covalent · polar covalent · nonpolar covalent
- E. ionic · ionic · polar covalent · nonpolar covalent

✓ Practice 5 answer: D



electronegativity: Li = 0.98 · H = 2.20 · N = 3.04 · F = 3.98

Li-F: **3.00** · H-F: **1.78** · N-F: **0.94** · F-F: **0**

→ ionic · polar covalent · polar covalent · nonpolar covalent → **answer D**

A called every large difference ionic: H-F (1.78) and N-F (0.94) join two nonmetals, which share. B called a real difference nonpolar: 0.94 is not zero. C took fluorine's own high electronegativity for polarity; in F-F, $3.98 - 3.98 = 0$. E read the 1.7 boundary as a wall: H-F is two nonmetals sharing a pair, polar covalent.

 **Sense check:** Only the difference counts, and at the edge the metal test decides. Li-F is the one metal-nonmetal pair. ✓



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?
3. Is the bonding in K_2SO_4 ionic, covalent, or both? Name which bonds are which.

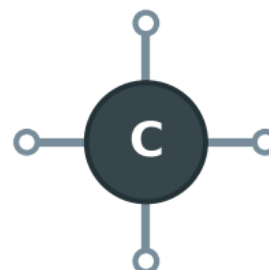
A bond type tells you how one shared pair of electrons is held. A whole molecule is built from several bonds arranged in space, and drawing that arrangement is the next step: Lewis structures.

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

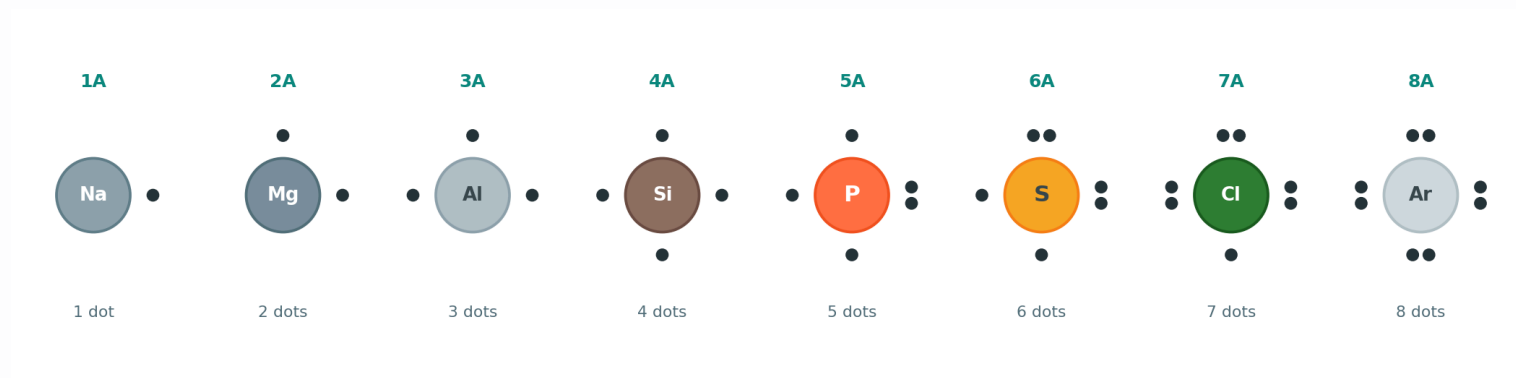
Only the outermost electrons, the valence electrons, form bonds. For a main-group atom their count is the ones digit of the group number.

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

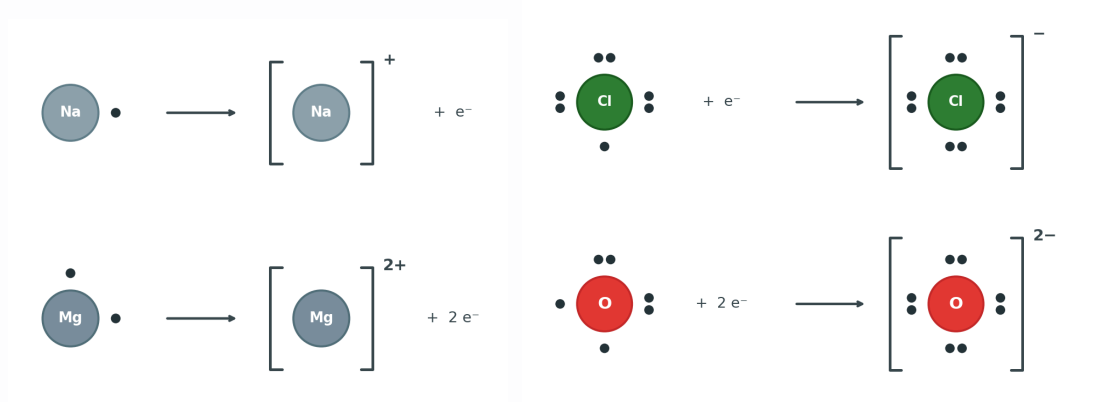
The Lewis symbol: valence electrons as dots

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.

Ionic bonds form by electron transfer

A metal atom gives its valence electrons to a nonmetal atom. Both ions end with a filled shell. The opposite charges attract, and that attraction holds the ions together as an ionic bond.



Na loses 1 e^- · Cl gains $1 \text{ e}^- \rightarrow \text{NaCl}$
 $1 \text{ e}^- \text{ lost} = 1 \text{ e}^- \text{ gained } \checkmark \cdot \text{charges: } (+1) + (-1) = 0 \checkmark$

Electrons lost equal electrons gained

Magnesium has two valence electrons to give. Each bromine needs only one to reach eight. Two bromine atoms take the pair, so the formula is MgBr_2 . Balancing the electrons balances the charges.



Mg loses 2 e^{-} · each Br gains $1 \text{ e}^{-} \rightarrow 1 \text{ Mg} : 2 \text{ Br} \rightarrow \text{MgBr}_2$

$2 \text{ e}^{-} \text{ lost} = 2 \times 1 \text{ e}^{-} \text{ gained} \checkmark$ · charges: $(+2) + 2(-1) = 0 \checkmark$

Your turn: lithium and oxygen

Li: group 1 · O: group 16

the metal gives electrons · the nonmetal takes them until it holds eight

Draw both Lewis symbols, arrow the transferred electrons, then fill the blanks and write the formula.

step	count
electrons each Li loses	_____
electrons one O gains	_____
Li atoms needed per O	_____

$\text{Li} \cdot \text{Li} \cdot + \text{O with 6 dots} \rightarrow 2 \text{Li}^+ + [\text{O with 8 dots}]^{2-} \rightarrow \text{Li}_2\text{O}$

each Li loses 1 · O gains 2 · 2 Li per O · charges: $2(+1) + (-2) = 0$ ✓



Practice 3: aluminum and sulfur

aluminum + sulfur $\rightarrow$ an ionic compound

Al: group 13 · S: group 16

What is the formula of the ionic compound that aluminum and sulfur form?

- A. AlS
- B. Al_2S_3
- C. Al_3S_2
- D. Al_6S_3

✓ Practice 3 answer: B

each Al loses 3 e^- · each S gains 2 e^- · $2 \times 3 = 3 \times 2 = 6 \rightarrow \text{Al}_2\text{S}_3 \rightarrow \text{answer B}$
 $6\text{ e}^- \text{ lost} = 6\text{ e}^- \text{ gained } \checkmark$ · charges: $2(+3) + 3(-2) = 0 \checkmark$

A paired one Al with one S without counting electrons: 3 lost against 2 gained. C wrote each element's own electron count as its own subscript: Al_3S_2 has $3 \times 3 = 9$ lost against $2 \times 2 = 4$ gained. D crossed the dot counts, 3 and 6, instead of the electrons that move: Al_6S_3 has $6 \times 3 = 18$ lost against $3 \times 2 = 6$ gained.

 **Sense check:** Neither 3 nor 2 divides the other, so both subscripts exceed 1. The electrons balance at 6, the smallest number both 3 and 2 divide into.

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

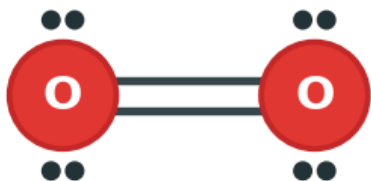
A shared pair drawn as a line is a bond: one pair single, two double, three triple. Each main-group atom aims for eight electrons, an octet. Hydrogen aims for two, a duet.

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

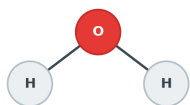
Step 2 • Draw the skeleton



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.

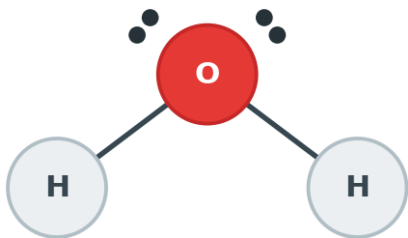
✓ Worked example 1: the oxygen lone pairs

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



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 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

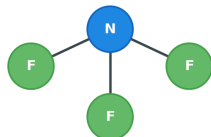
Step 1 • Count the valence electrons

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.

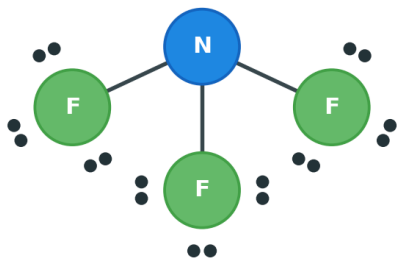
✓ Worked example 2: the fluorine octets

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.

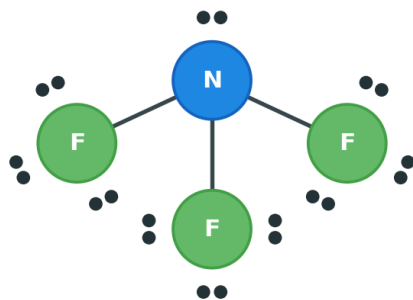
✓ Worked example 2: the nitrogen lone pair

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

NF_3 : $5 + 3(7) = 26 \checkmark$ · $5 + 7 = 12 \times$
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

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

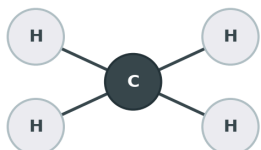
1 C (group 14 $\rightarrow$ 4) $\cdot$ 4 H (group 1 $\rightarrow$ 1) $\cdot$ 8 valence electrons

Carbon is the central atom. Fill the blanks.

step	count
electrons in the 4 C-H bonds	_____
electrons left for lone pairs	_____

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

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 4: 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

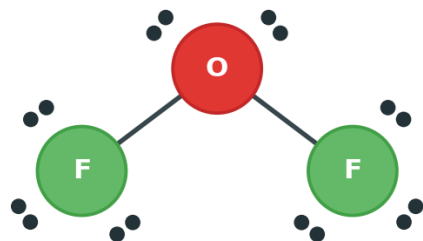
✓ Practice 4 answer: D

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

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: $20 - 2 = 18$.

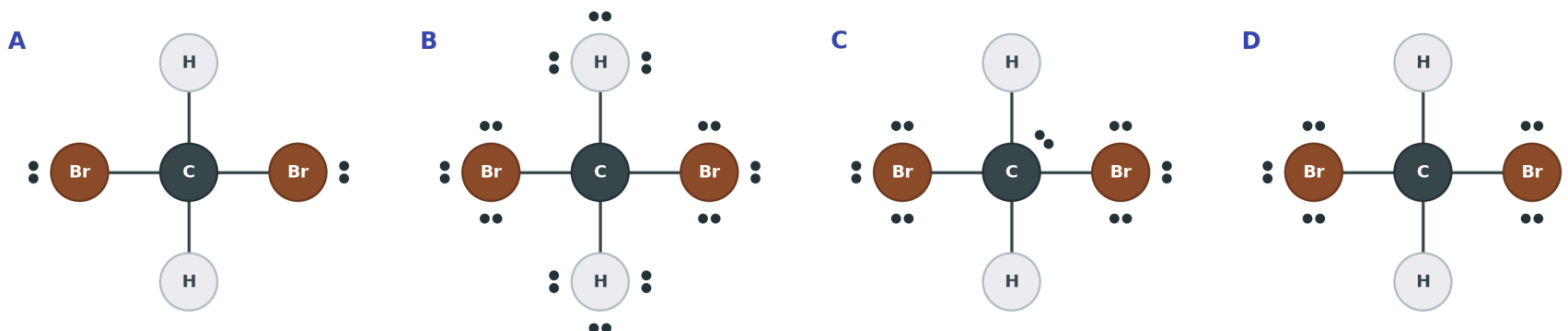
🧠 **Sense check:** 20 valence electrons place as 2 O-F bonds (4) and 8 lone pairs (16): $4 + 16 = 20$, every atom an octet.



🎯 Practice 5: the structure of CH_2Br_2



wanted: the correct Lewis structure



Which drawing is the correct Lewis structure of CH_2Br_2 ?

A. one lone pair on each Br

B. three lone pairs on each Br and on each H

C. three lone pairs on each Br, one on C

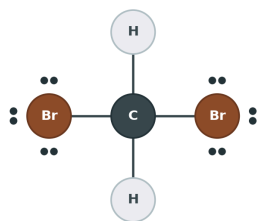
D. three lone pairs on each Br, none on C or H

✓ Practice 5 answer: D

$$\text{CH}_2\text{Br}_2: 4 + 2(1) + 2(7) = 20 \rightarrow \text{answer D}$$

4 bonds (8) + 3 lone pairs on each Br (12) · 20 – 8 – 12 = 0 left for C · C 8 ✓ · each Br 8 ✓ · each H a duet ✓

A counted each element once: $4 + 1 + 7 = 12$, leaving $12 - 8 = 4$ after the bonds, one pair per Br. B gave hydrogen an octet: $8 + 12 + 12 = 32$ drawn from a pool of 20; H takes a duet. C put a lone pair on carbon with none left: $8 + 12 + 2 = 22$, and carbon holds 10.



 **Sense check:** Carbon's four bonds already give it eight. With nothing left over, the center carries no lone pair, as in CH_4 .



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

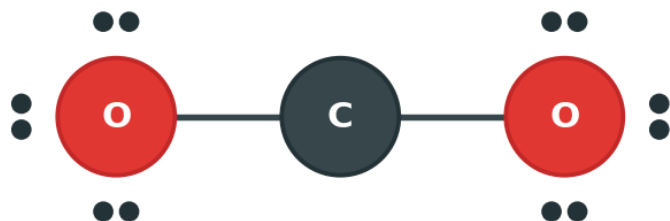


🧠 **Sense check:** Four of the 16 electrons fill the two C-O bonds. Twelve remain for the outer octets.

✓ 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.

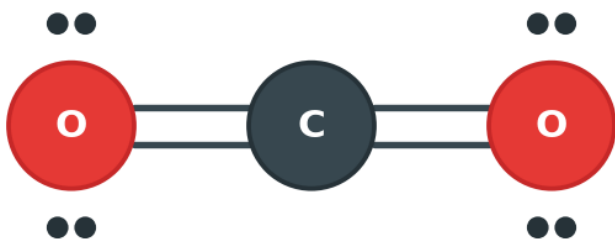
✓ Worked example 3: two double bonds

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

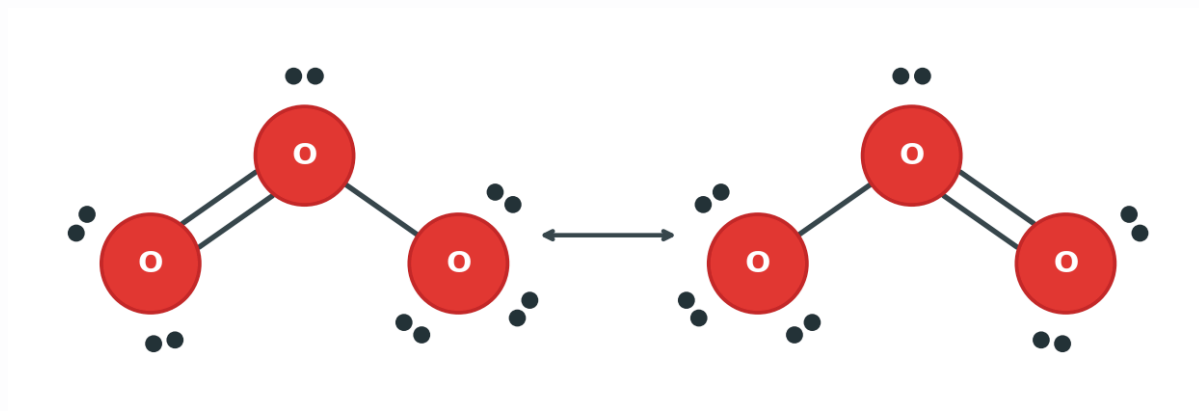


■ in bonds
 ■ lone pairs on outer atoms
 ■ lone pairs on the center
 ■ lone pairs moved into bonds

 **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. ✓

Resonance: two correct drawings of ozone

Ozone's 18 valence electrons need one O=O double bond, and it fits on either side. Both drawings are correct. The measured bonds are identical: the real molecule, the resonance hybrid, blends the two.



measured: both O-O bonds the same length, between a single and a double bond



Worked example 4: the carbonate ion

Step 1 • Count the valence electrons

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

1 C (group 14 $\rightarrow$ 4) $\cdot$ 3 O (group 16 $\rightarrow$ 6) $\cdot$ the 2– charge adds 2 $\cdot$ 24 to place

Carbon takes the center. Draw the Lewis structure, then count its equivalent drawings.

✓ Worked example 4: solution

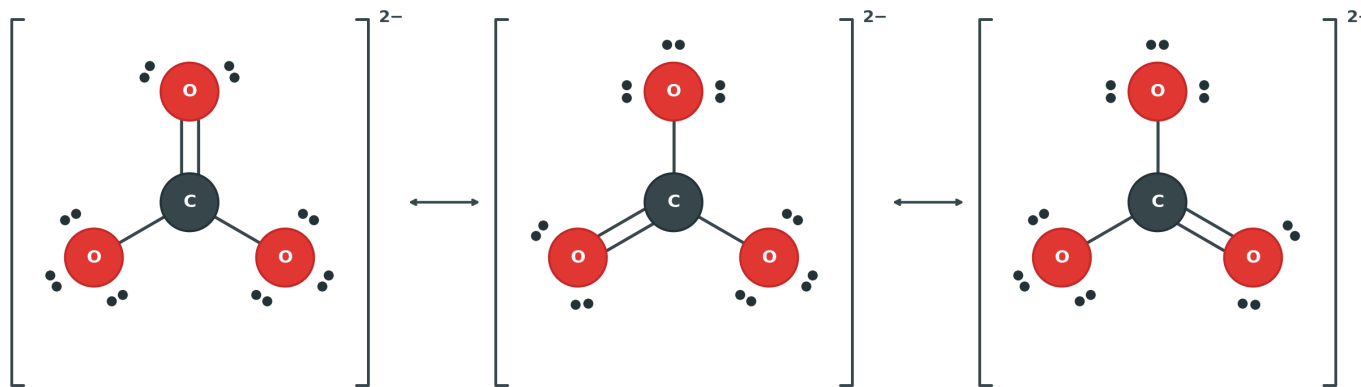
Step 2 · Draw the skeleton

Step 3 · Complete the outer octets

Step 4 · Place leftovers on the center

3 C-O bonds (6) + 3 lone pairs on each O (18) = 24 · 24 – 6 – 18 = 0 left
all 24 placed · carbon holds only 6, one pair short of an octet

Step 5 · Form multiple bonds



 **Sense check:** Any of the three O atoms can hold the C=O: three equivalent drawings. The real ion has three identical C-O bonds, between single and double.

✓ Worked example 4: the electron budget

CO_3^{2-} : 24 valence electrons = 8 in bonds + 16 in lone pairs on O
all five steps used · step 4 adds nothing: none left

CO_3^{2-} : 24 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 same route as CO_2 , but only 1 lone pair moves: $6 + 2 = 8$ in bonds. It can come from any of the 3 O atoms. ✓

When resonance appears

Resonance appears when a multiple bond can sit in more than one place and every atom keeps its octet. Only bonds and lone pairs shift; atoms never move. The drawings are usually equivalent, but need not be.

CO_3^{2-} , NO_3^- : 3 structures · O_3 , NO_2^- : 2 structures · H_2O , C_2H_4 : 1 structure
equivalent drawings in each · H_2O has no multiple bond · the $\text{C}=\text{C}$ in C_2H_4 has one place

SCN^- : $\text{S}=\text{C}=\text{N} \leftrightarrow \text{S}\equiv\text{C}-\text{N} \leftrightarrow \text{S}-\text{C}\equiv\text{N}$: 3 structures, not alike
16 valence electrons · every atom has an octet in all three · the bonds shift between different atoms

Practice 6: hydrogen cyanide

HCN
skeleton H-C-N

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

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

✓ Practice 6 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.

Practice 7: an octet check on N₂O



skeleton N-N-O · lone pairs placed to use the valence electrons

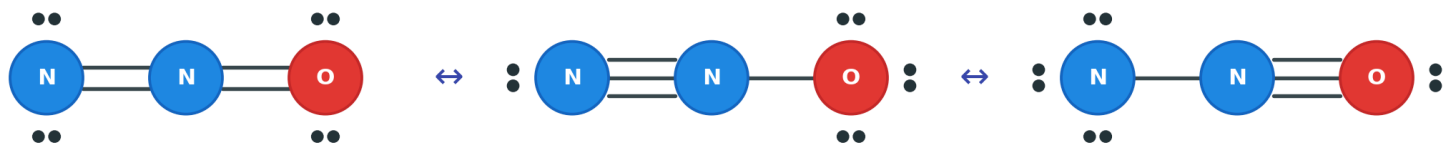
Which drawing of N₂O breaks the octet rule?

- A. N=N=O
- B. N≡N-O
- C. N≡N=O
- D. N-N≡O
- E. none: all four are valid

✓ Practice 7 answer: C

N_2O : $2(5) + 6 = 16$ · $\text{N}\equiv\text{N}=\text{O}$: 5 bond pairs = 10 shared · $16 - 10 = 6$ in lone pairs →
answer C

center N: triple (6) + double (4) = 10, past an octet X · the other three: every atom 8 ✓



A counted two double bonds as too many: the center N has $4 + 4 = 8$. B took the single-bonded O as unfinished: $2 + 6 = 8$ with three lone pairs. D took O's triple bond as too many: $6 + 2 = 8$ with one lone pair. E stopped before the center atom, where $\text{N}\equiv\text{N}=\text{O}$ puts 10.

 **Sense check:** Count the center atom first: a triple and a double bond on one atom already make 10. The three valid drawings are the resonance structures of N_2O . ✓



Check yourself

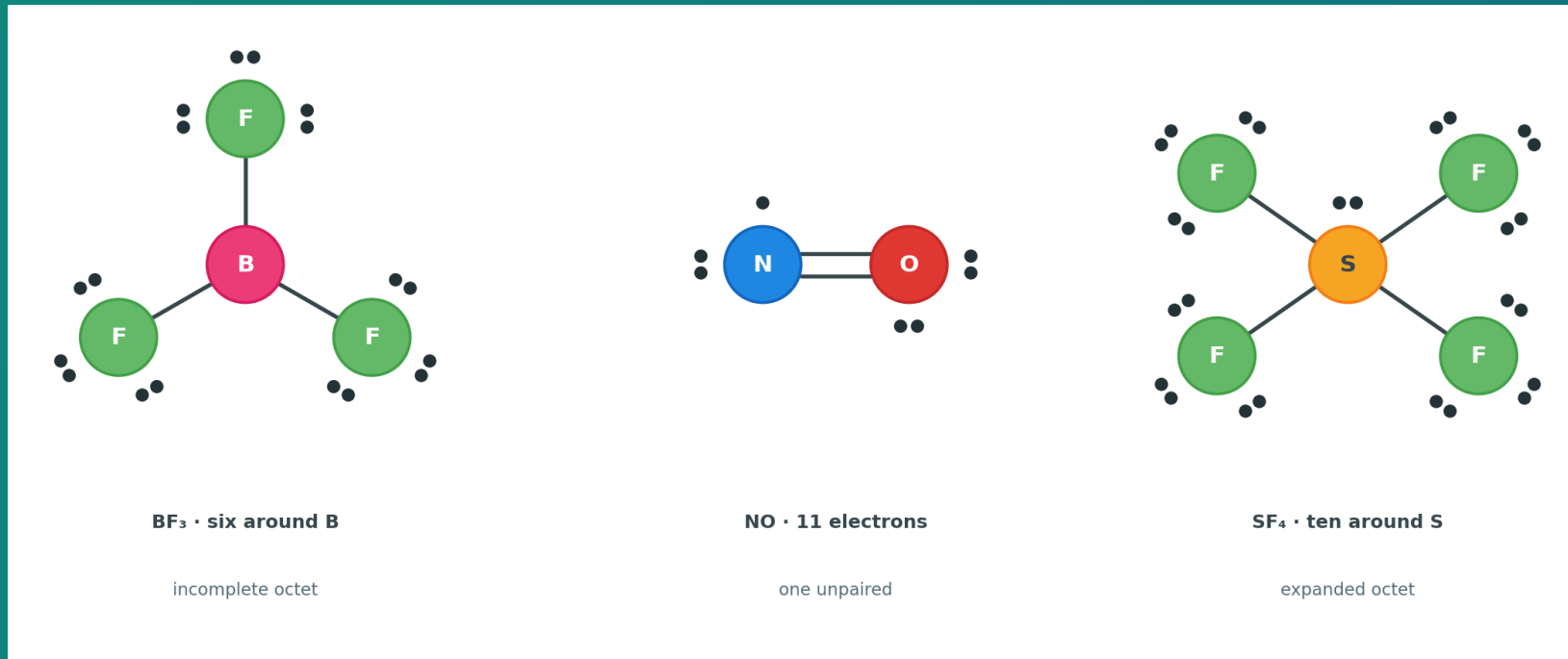
1. Count the valence electrons in NH_3 , and name the central atom.
2. In CO_2 , why do single bonds fail and double bonds succeed?

The octet rule builds most structures, not all: boron can settle at six electrons, and heavier central atoms can hold more than eight.

3 · 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.

BF_3 : six around B · NO: 11 electrons, one unpaired · SF_4 : 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: $5 \times 6 = 30$

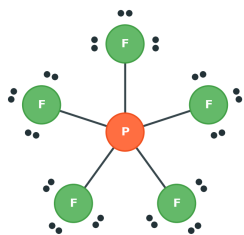
$10 + 30 = 40$ · every electron placed · every F has its octet

🧠 **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.

⚠ 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.

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: naming the exception

	molecule
1	BCl_3
2	ClO_2
3	PCl_5

Name the octet exception each molecule shows, in order.

- A. expanded · odd-electron · expanded
- B. odd-electron · odd-electron · expanded
- C. incomplete · odd-electron · expanded
- D. incomplete · incomplete · expanded

✓ Practice 1 answer: C

BCl_3 : $3 + 3(7) = 24$, B at 6 · ClO_2 : $7 + 2(6) = 19$, odd · PCl_5 : 5 bonds, P at 10
incomplete · odd-electron · expanded → answer C

A read the period of the chlorines instead of the center. Boron is in period 2 and stops at six. B took boron's own count, 3, as the molecule's; the total, 24, is even. D called ClO_2 incomplete because one atom falls short of eight; an odd total, 19, puts it in the odd-electron family.

🧠 **Sense check:** Boron center: incomplete. Odd total: odd-electron. Five bonds on period-3 phosphorus: expanded.

✓



Practice 2: 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 2 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.

Practice 3: an unpaired electron

Which species carries an unpaired electron?

- A. NO_2^-
- B. NO_2
- C. N_2O_4
- D. BF_3

✓ Practice 3 answer: B

NO_2 : $5 + 2(6) = 17$, odd $\rightarrow$ one electron stays unpaired $\rightarrow$ answer B

NO_2^- : $5 + 2(6) + 1 = 18$ · N_2O_4 : $2(5) + 4(6) = 34$ · BF_3 : $3 + 3(7) = 24$ · all even

A dropped the 1– charge: $5 + 2(6) = 17$ looks odd, but the added electron makes 18, and 18 pairs fully. C judged by nitrogen's own count, 5; two nitrogens give $2(5) + 4(6) = 34$, even. D took boron's short octet for an unpaired electron; BF_3 holds 24 electrons, all in pairs.

 **Sense check:** Only an odd total leaves a single electron. Two NO_2 molecules pair their odd electrons to form N_2O_4 .

✓



Worked example 2: boron trifluoride

Step 1 • Count the valence electrons

$$\text{BF}_3: 3 + 3(7) = 24$$

1 B (group 13 $\rightarrow$ 3) $\cdot$ 3 F (group 17 $\rightarrow$ 7) $\cdot$ 24 to place

Three fluorines bond to one boron. Draw the Lewis structure and count the electrons around boron.

✓ Worked example 2: solution

BF_3 : $3 + 3(7) = 24$ valence electrons · skeleton: three B-F single bonds

Step 2 · Build the skeleton, then the outer octets

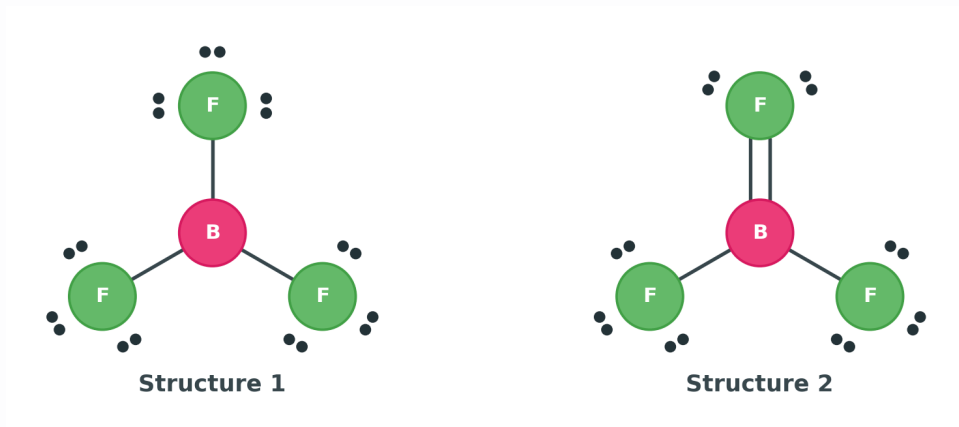
3 bonds: $3 \times 2 = 6$ · 3 F $\times$ 3 lone pairs: $3 \times 6 = 18$
 $6 + 18 = 24$ · every electron placed · every F has its octet

Step 3 · Leftovers go on the central atom

$24 - 6 - 18 = 0$ leftover · B: 3 bonds = 6 electrons
 nothing is left for boron · B stops at six

🧠 **Sense check:** All 24 electrons are placed, every fluorine is full, and boron holds six: an incomplete octet.

✓ Worked example 2: leave boron at six



Structure 1: B at 6 · Structure 2: B at 8 through one B=F double bond
both place all 24 · every F at 8 in both

The tempting fix Pull a fluorine lone pair into a B=F double bond (Structure 2). Reject it: boron, like beryllium, accepts fewer than eight.

 **Sense check:** Structure 1 is BF_3 : three single bonds, every F at eight, B at six. ✓

Take-home: boron and beryllium stop short

If nothing is left and the center is boron or beryllium, stop at fewer than eight. Never pull an outer lone pair into a double bond just to complete its octet.

BF_3 : 24 electrons, all paired · every F at 8 · B at 6

incomplete octet: the center falls short, the total is even · never force a B or Be octet

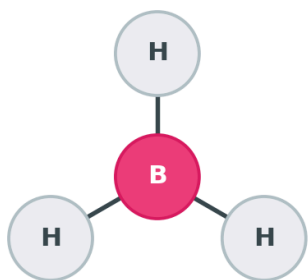
Your turn: borane, BH_3

BH_3 : $3 + 3(1) = 6$ valence electrons

skeleton: three B-H single bonds · each H is full with one bond, 2 electrons

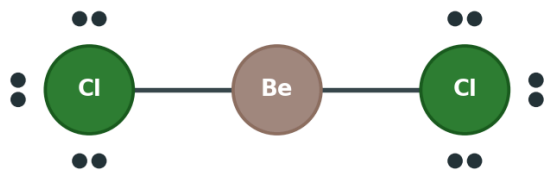
$3 \times 2 = 6$ · $6 - 6 = 0$ leftover · B: 3 bonds = 6 electrons

incomplete octet on boron · nothing is left, and boron accepts six: the same call as BF_3

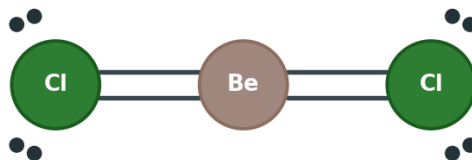


Practice 4: beryllium chloride

BeCl_2 : $2 + 2(7) = 16$ valence electrons · both structures place all 16



Structure 1

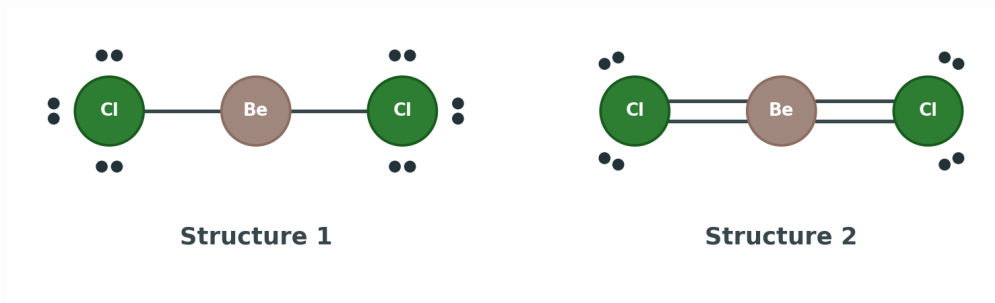


Structure 2

Which structure is best, and how many electrons surround beryllium in it?

- A. Structure 1, 4 electrons
- B. Structure 2, 8 electrons
- C. Structure 1, 8 electrons
- D. Structure 2, 4 electrons

✓ Practice 4 answer: A



structure 1: Be has two single bonds: $2 + 2 = 4$ electrons $\rightarrow$ answer A
 each Cl keeps three lone pairs and one bond: $8 \checkmark$ · Be, like B, accepts fewer than eight

B forced the octet: pulling a Cl lone pair into each bond gives Be $4 + 4 = 8$, the same move BF_3 rejected. C picked the right structure but assumed eight: two single bonds hold only $2 + 2 = 4$. D counted each double bond as one pair: two double bonds put $4 + 4 = 8$ around Be.

 **Sense check:** Beryllium, like boron, accepts fewer than eight. Cl-Be-Cl with single bonds: four electrons around Be. ✓



Check yourself

1. Which periods allow an expanded octet, and which four elements never expand?
2. How many electrons surround boron in BF_3 , and why is that structure left as an incomplete octet?

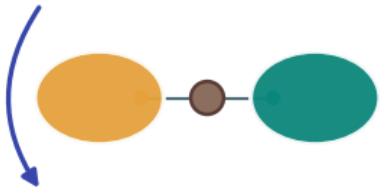
The count of bonds and lone pairs around a central atom, octet or expanded, is exactly what fixes a molecule's three-dimensional shape.

4 · VSEPR & Molecular Shape

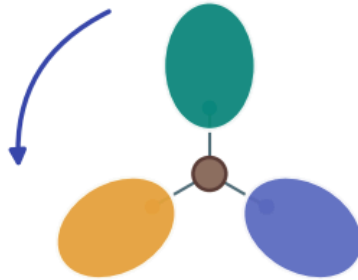
Predict a molecule's electron-pair geometry and its shape by counting the electron domains on the central atom and reading which corners lone pairs take.

What sets a molecule's shape

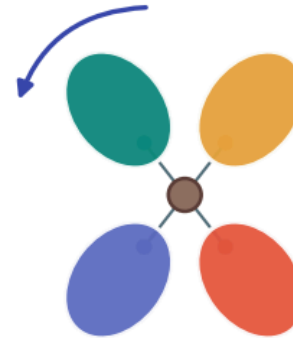
Tied at one knot, balloons crowd, then spread as far apart as they can.



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)

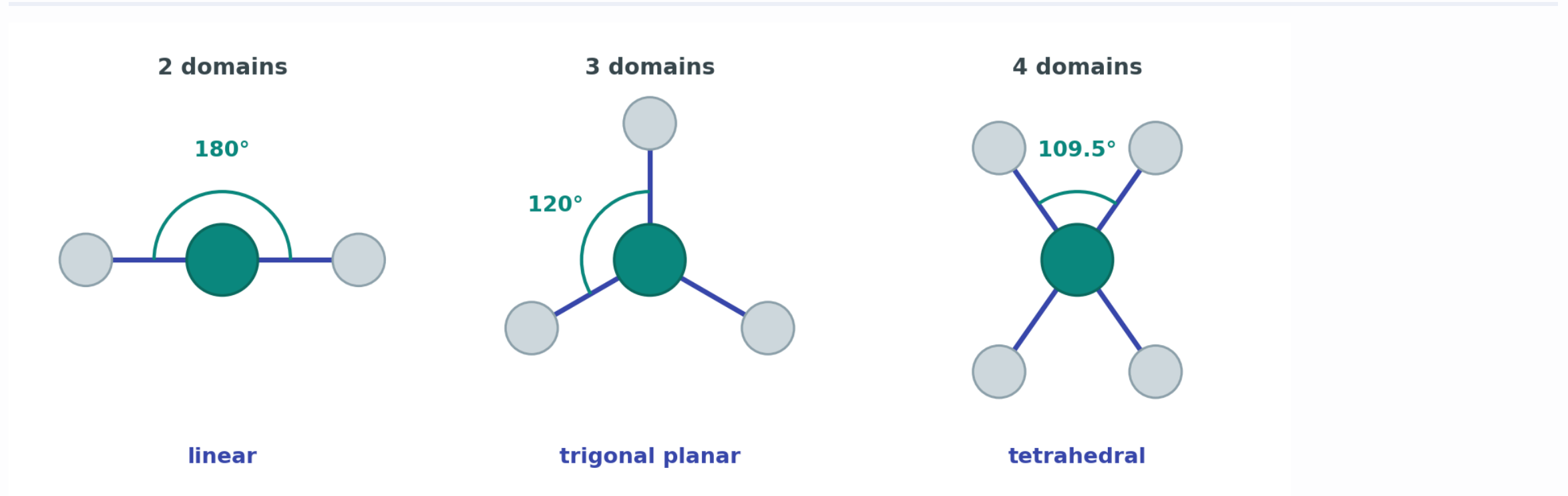
NH₃: N bonded to 3 H, plus 1 lone pair

3 bonded atoms + 1 lone pair → 4 domains

$$\text{domains} = X + E$$

X = atoms bonded to the central atom · E = lone pairs on the central atom

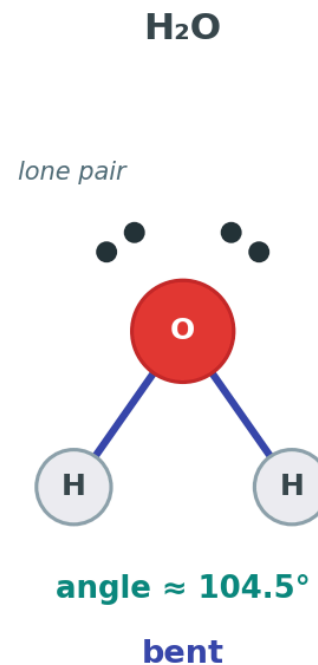
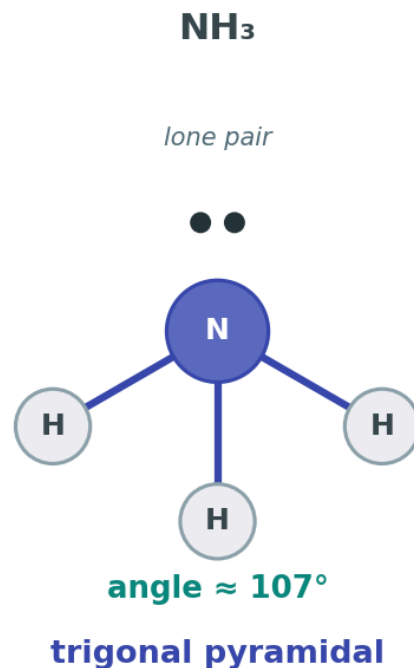
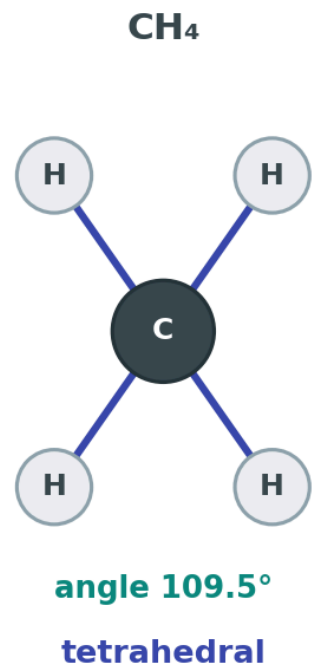
From domain count to electron-pair geometry



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

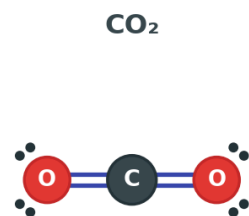
Same electron arrangement, 4 tetrahedral domains: three different shapes



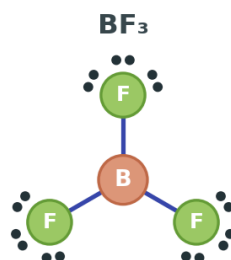
Lone pairs take up domains but hold no atom. The molecular shape names only where the atoms sit. Four tetrahedral domains read as tetrahedral, trigonal pyramidal, or bent as lone pairs replace atoms.

The common shapes side by side

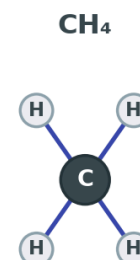
The lone pairs are the dark dots: they claim a domain but hold no atom, so the shape reads only where the atoms sit.



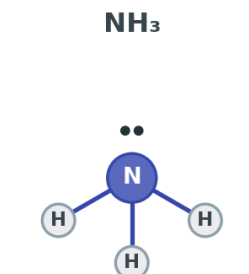
linear
AX₂ · 180°



trigonal planar
AX₃ · 120°



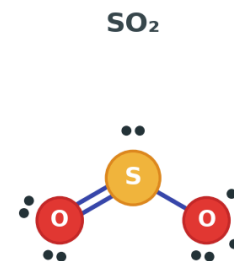
tetrahedral
AX₄ · 109.5°



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.



Guided example: oxygen difluoride



only the formula · no bonds or lone pairs drawn yet

A formula does not show the lone pairs. Draw the Lewis structure first, then read X and E on oxygen.

Find the molecular shape of OF_2 .

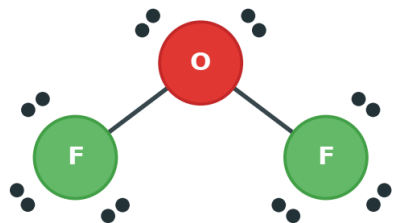
✓ Guided example: the Lewis structure

First · Draw the Lewis structure

OF_2 : $6 + 2(7) = 20$ valence electrons

F-O-F: 2 bonds use 4 · each F takes 6 more: $2 \times 6 = 12$ · $20 - 4 - 12 = 4$ left on O

Read X and E



Two fluorines bond to oxygen: $X = 2$. Two lone pairs sit on oxygen: $E = 2$.

 **Sense check:** Every atom has an octet, and $4 + 16 = 20$ electrons are placed. ✓

✓ Guided example: the shape

OF_2 : F-O-F with 2 lone pairs on O

from the Lewis structure: X = 2 bonded F · E = 2 lone pairs on O

Step 1 · Count the electron domains

Step 2 · Name the electron-pair geometry

Step 3 · Name the molecular shape

$\text{OF}_2 \rightarrow$ bent

4 domains · tetrahedral geometry · 2 corners hold lone pairs

 **Sense check:** The formula alone hid the two lone pairs. The Lewis structure found them, and they bend the molecule.



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 → 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.



Practice 2



phosphorus trichloride, a starting material for weed killers

Which is the molecular shape of PCl_3 ?

- A. Tetrahedral: four electron domains around phosphorus
- B. Trigonal pyramidal: three chlorines and one lone pair on phosphorus
- C. Trigonal planar: three chlorines and no lone pair on phosphorus
- D. Bent: three domains, one of them a lone pair

✓ Practice 2 answer: B

PCl_3 : $5 + 3(7) = 26$ · 3 bonds use $3 \times 2 = 6$ · each Cl takes 6: $3 \times 6 = 18$ · $26 - 6 - 18 = 2$ left → 1 lone pair on P

$X = 3$ · $E = 1$ · $3 + 1 = 4$ domains · tetrahedral geometry · AX_3E , trigonal pyramidal → answer B

A named the electron-pair geometry; one corner holds the lone pair, not an atom. C stopped after the outer octets: $26 - 6 - 18 = 2$ electrons remain, and they sit on phosphorus. D counted three domains with one a lone pair, which leaves room for only two chlorines. PCl_3 has three bonded Cl plus the lone pair: $3 + 1 = 4$ domains.

 **Sense check:** The formula showed three chlorines. The Lewis structure showed the lone pair that makes the pyramid. ✓



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 → AX_2E
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 3



the nitrite ion, used to cure bacon and ham

What are the shape and the bond angle of the nitrite ion?

- A. Tetrahedral, 109.5°
- B. Trigonal planar, 120°
- C. Bent, $<109.5^\circ$
- D. Bent, $<120^\circ$
- E. Linear, 180°

✓ Practice 3 answer: D

NO_2^- : $5 + 2(6) + 1 = 18$ · 2 bonds use 4 · each O takes 6: $2 \times 6 = 12$ · $18 - 4 - 12 = 2$ left → 1 lone pair on N

N has only $4 + 2 = 6 \rightarrow$ one O shares a pair: $\text{O}=\text{N}-\text{O}^-$, 2 resonance forms · $X = 2$ · $E = 1$ either way · 3 domains · AX_2E , bent, just under $120^\circ \rightarrow$ answer D

B named the electron-pair geometry; the lone pair takes a corner, so the atoms are bent. C counted the $\text{N}=\text{O}$ double bond as two domains: $2 + 1 + 1 = 4$, a tetrahedral base. A made the same count and then named that geometry. E subtracted the charge: $5 + 2(6) - 1 = 16$ leaves $16 - 4 - 12 = 0$ for N, so $\text{O}=\text{N}=\text{O}$ would be linear.

🧠 **Sense check:** The 1– charge adds the electron that finishes nitrogen's lone pair. The $\text{N}=\text{O}$ double bond still counts once. ✓



Practice 4

ClO_3^- : Cl bonded to 3 O

a student draws three Cl-O bonds, no lone pair on Cl, and names the shape trigonal planar

The chlorate ion, ClO_3^- , is found in match heads. Evaluate the work.

- A. Correct: three bonded atoms and no lone pair give trigonal planar, 120°
- B. Wrong: chlorine keeps one lone pair, so four domains make the shape tetrahedral
- C. Wrong: chlorine keeps one lone pair, so four domains make the shape trigonal pyramidal
- D. Wrong: chlorine keeps one lone pair, so four domains make the shape bent

✓ Practice 4 answer: C

ClO_3^- : $7 + 3(6) + 1 = 26$ · 3 bonds use 6 · each O takes 6: $3 \times 6 = 18$ · $26 - 6 - 18 = 2$ left → 1 lone pair on Cl

Cl: $6 + 2 = 8$ ✓ · $X = 3$ · $E = 1$ · $3 + 1 = 4$ domains · tetrahedral geometry · AX_3E , trigonal pyramidal, just under 109.5° → answer C

A dropped the leftover pair: the charge's electron makes 26, and 2 of them stay on chlorine. B named the electron-pair geometry; one corner holds the lone pair, not an atom. D took the water row: bent needs two lone pairs and two atoms, and ClO_3^- has three bonded O and one lone pair.

🧠 **Sense check:** Same row as NH_3 and PCl_3 : three atoms, one lone pair. Count the charge before reading the shape.

✓

Check yourself

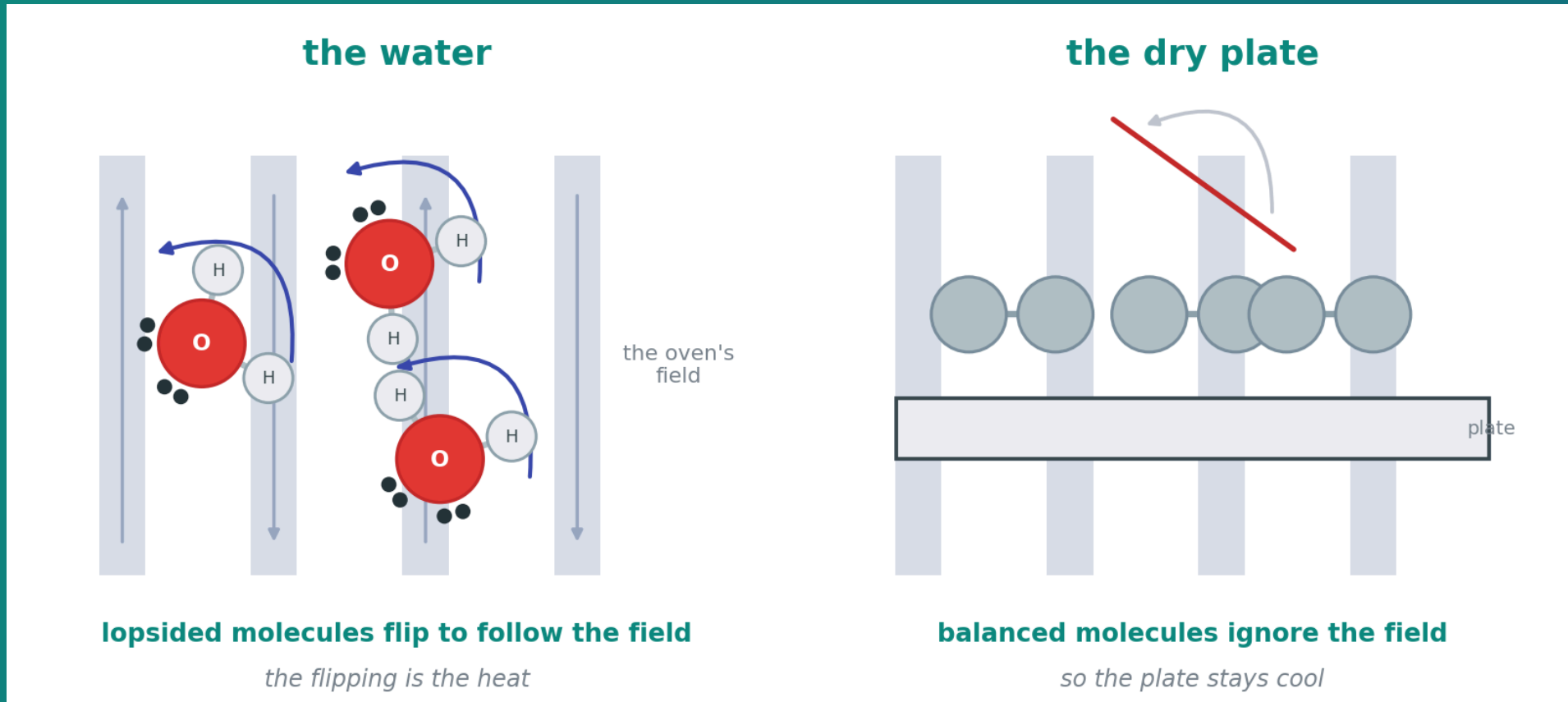
1. Draw the Lewis structure of PH_3 . Count X and E on phosphorus, then name the electron-pair geometry and the shape.
2. Why do CO_2 and SO_2 have different shapes? Draw both Lewis structures and name each one's AXE row.

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.

5 • 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

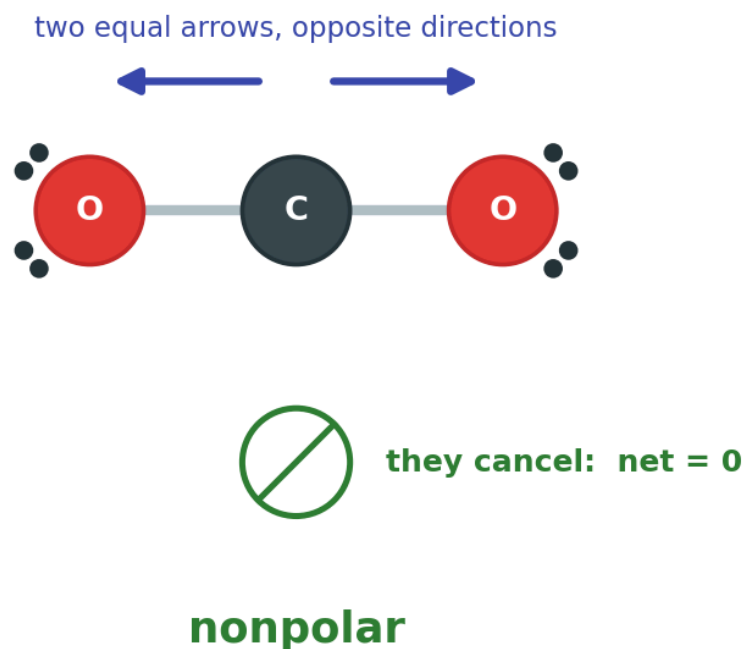


the shared electrons sit closer to chlorine

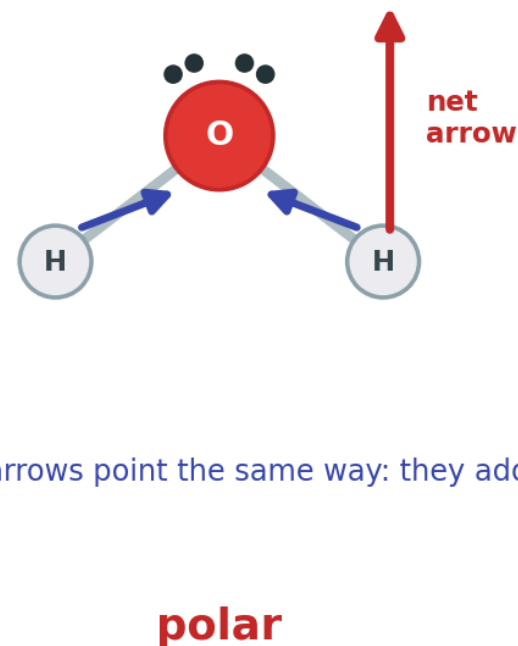
Draw each polar bond as a dipole arrow. It points to the more electronegative atom, the end that pulls the shared electrons closer. A bigger electronegativity difference means a stronger pull.

When arrows cancel, and when they add

CO₂ · linear

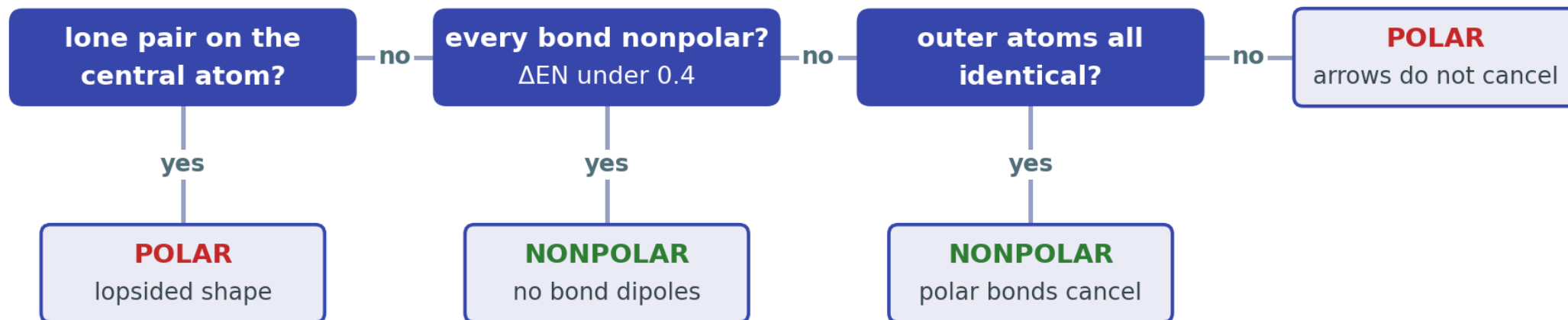


H₂O · bent



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.

When a molecule is nonpolar



A molecule with a lone pair on its central atom is polar. Otherwise it is nonpolar when every bond is nonpolar, or when its polar bonds cancel (identical outer atoms).



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

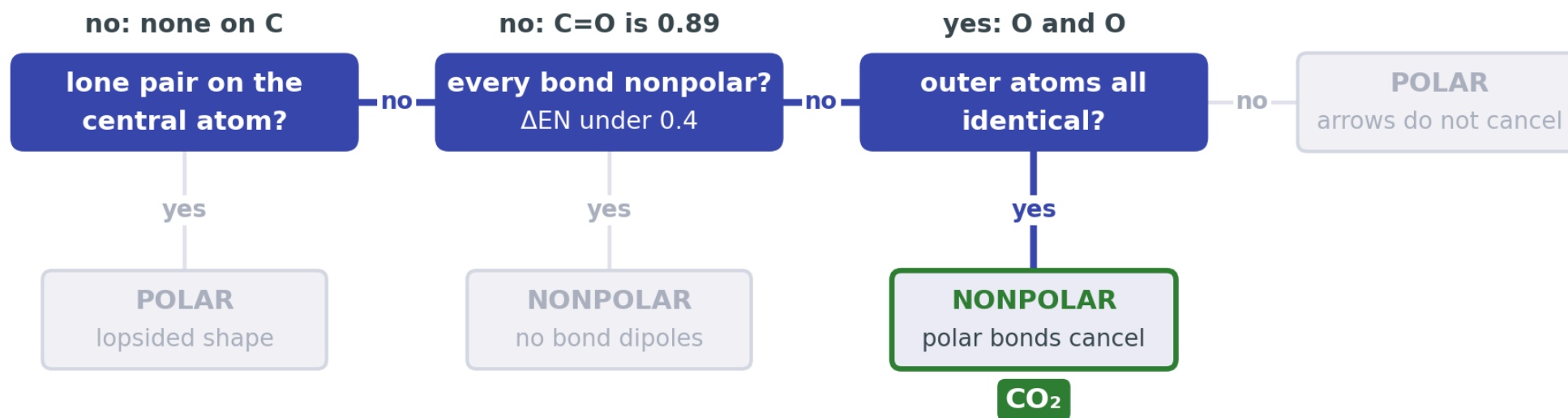
0.89 toward one O – 0.89 toward the other O = **0 net**

 **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. ✓

Practice 2: carbon disulfide

CS₂: carbon disulfide

electronegativity: C = 2.55 · S = 2.58 · wanted: polar or nonpolar, and why

Is carbon disulfide polar or nonpolar?

- A. Nonpolar: C and S pull almost equally, and the linear shape would cancel any dipoles
- B. Polar: S out-pulls C, so each bond is polar and so is the molecule
- C. Polar: a lone pair on carbon bends the molecule
- D. Polar: the lone pairs on the sulfur atoms bend the molecule

✓ Practice 2 answer: A



two C=S bonds use 8 · four lone pairs on the S atoms use 8 · none left on C · linear

$\Delta\text{EN}(\text{C}=\text{S}): 2.58 - 2.55 = \mathbf{0.03}$ · under 0.4: nonpolar bonds → nonpolar → **answer A**

B skipped the difference: 0.03 is far under 0.4, and even polar bonds would cancel in a straight line. C gave carbon a lone pair; its four valence electrons all go into the two double bonds. D bent the molecule with the outer atoms' lone pairs; only lone pairs on the central atom set the shape.

 **Sense check:** Nonpolar twice over: the bonds are nonpolar, and the linear shape is symmetric. ✓

Practice 3: hydrogen sulfide

H₂S: hydrogen sulfide

electronegativity: H = 2.20 · S = 2.58 · wanted: polar or nonpolar, and why

Is hydrogen sulfide polar or nonpolar?

- A. Nonpolar: the molecule is linear, so the two S-H arrows cancel
- B. Nonpolar: $\Delta\text{EN}(\text{S-H})$ is under 0.4, so the bonds are nonpolar
- C. Nonpolar: sulfur's two lone pairs sit on opposite sides and balance each other
- D. Polar: two lone pairs on sulfur make the molecule bent and lopsided

✓ Practice 3 answer: D



two S-H bonds use 4 · 8 – 4 = 4 left = 2 lone pairs on S · 4 domains · bent

lone pair on the central S? yes → **polar** → **answer D**

A missed the two lone pairs and drew H₂S straight, like CS₂. B skipped the first question: 2.58 – 2.20 = 0.38 is under 0.4, but a central lone pair makes the molecule polar anyway. C read the flat drawing as the shape; four domains spread in three dimensions, and both lone pairs sit on one side of sulfur.

🧠 **Sense check:** Two outer atoms in CS₂ and H₂S, opposite answers: the lone pairs on sulfur make the difference. ✓



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

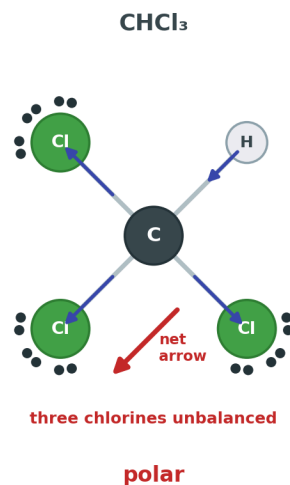
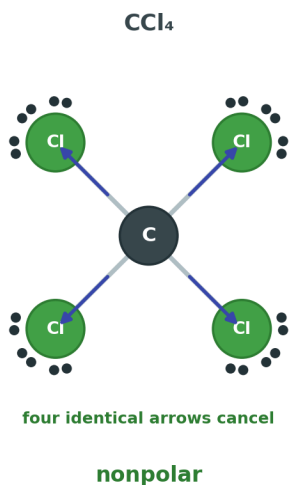
 **Sense check:** The C-Cl bonds are polar. Whether the molecule is polar now depends only on how the four arrows are arranged.

✓ 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

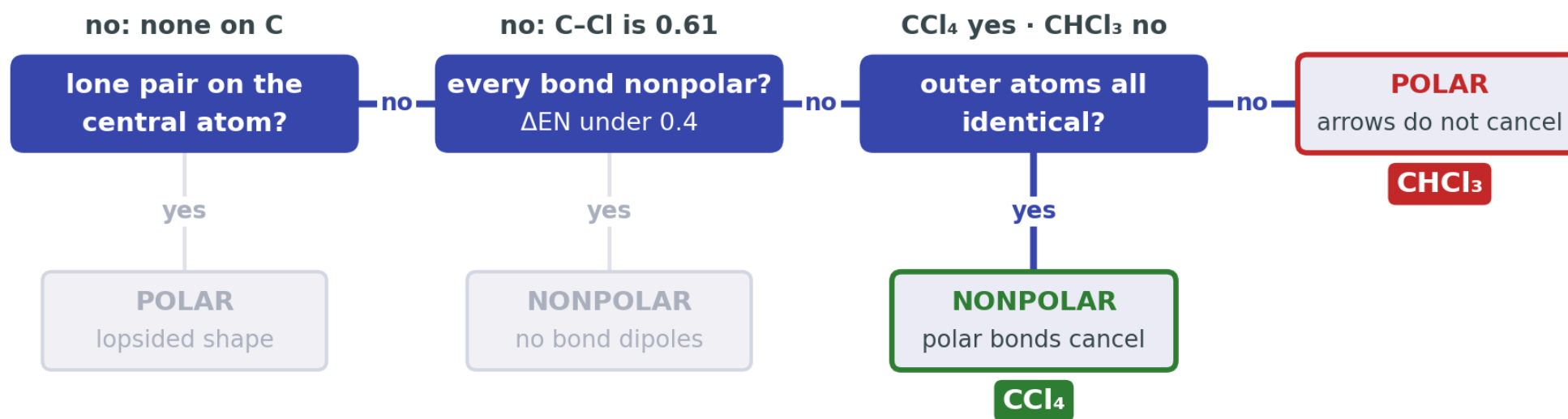


🧠 **Sense check:** CCl_4 's four identical arrows cancel; swapping one chlorine for hydrogen leaves the three chlorine arrows unbalanced, so a net arrow remains: polar. ✓

✓ 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 4: dichloromethane

CH₂Cl₂: dichloromethane

electronegativity: H = 2.20 · C = 2.55 · Cl = 3.16 · wanted: polar or nonpolar, and why

Is dichloromethane polar or nonpolar?

- A. Nonpolar: it is tetrahedral with no lone pair on carbon, so the arrows cancel
- B. Nonpolar: the two C-Cl arrows point opposite ways, as drawn on paper, and cancel
- C. Polar: the outer atoms are not identical, so the arrows do not cancel
- D. Polar: its bonds are polar, and polar bonds always make a polar molecule

✓ Practice 4 answer: C

CH₂Cl₂: tetrahedral, no lone pair on C

outer atoms: 2 H and 2 Cl, not identical · $\Delta\text{EN}(\text{C-Cl}) = 0.61$ · $\Delta\text{EN}(\text{C-H}) = 0.35$

two strong C-Cl arrows + two weak C-H → unbalanced: polar → answer C

A stopped at the lone-pair question; a tetrahedral shape cancels only when all four outer atoms are identical, as in CCl₄. B trusted the flat drawing; in three dimensions the two C-Cl bonds sit 109.5° apart, never opposite, so their arrows add. D used a rule CCl₄ breaks: four polar bonds, nonpolar molecule.

 **Sense check:** Swap outer atoms on a symmetric shape and the balance breaks: polar. ✓



Practice 5



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 5 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. Polarity also decides what mixes: polar water dissolves polar substances and leaves nonpolar oil apart. That rule, like dissolves like, returns with liquids and solutions.

Can you...?

- ☐ classify a compound's bonding as ionic, covalent, or both, and a bond as nonpolar covalent, polar covalent, or ionic from the electronegativity difference?
- ☐ write Lewis symbols for atoms and ions and show the electron transfer that gives an ionic compound its formula?
- ☐ draw Lewis structures for molecules and polyatomic ions, including multiple bonds, resonance forms, and the common octet exceptions?
- ☐ use VSEPR to predict the shape and bond angle of a molecule with up to four electron domains?
- ☐ combine bond polarity with shape to decide whether a molecule is polar?

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