

Lewis structures, shapes and hybridization of molecules

AP Chemistry · Worksheet answers

Each answer gives the electron count, the structure in words and the result in bold. Always check that the electrons drawn equal the electrons available and that formal charges add to the overall charge.

Level: AP Chemistry · **Curriculum:** AP Chem Unit 2

Problem 1 Classify the bonds

1. NaCl: $3.16 - 0.93 = 2.23$; HCl: $3.16 - 2.20 = 0.96$; Cl₂: 0.
2. **NaCl ionic, HCl polar covalent, Cl₂ nonpolar covalent.**
3. H is δ^+ and Cl is δ^- : $\text{H}^{\delta+}-\text{Cl}^{\delta-}$.

Problem 2 Count valence electrons

1. $2(1) + 6 = \mathbf{8}$.
2. $4 + 2(6) = \mathbf{16}$.
3. $5 + 4(1) - 1 = \mathbf{8}$.
4. $5 + 3(6) + 1 = \mathbf{24}$.

Problem 3 Simple Lewis structures

1. Carbon at the centre with four single bonds to H, no lone pairs: **8 electrons around C.**
2. $\text{H}-\text{C}\equiv\text{N}$ with one lone pair on N. Electron count: 2 (C-H bond) + 6 (triple bond) + 2 (lone pair) = 10, the same as the valence electrons available.
3. **One lone pair** on N.

Problem 4 Formal charges on ions

1. $5 - 0 - 4 = \mathbf{+1}$.
2. $6 - 6 - 1 = \mathbf{-1}$.

Problem 5 Water's shape

1. **4 domains** (2 bonds + 2 lone pairs).
2. The electron geometry is tetrahedral and the molecular shape is **bent**.
3. About **104.5°**.

Problem 6 Bond length and energy

1. Bond length **decreases** from 154 to 134 to 120 pm.
2. Bond energy **increases** from 347 to 614 to 839 kJ/mol.
3. More shared electron pairs pull the two nuclei closer together and hold them more tightly, so the bond is shorter and stronger.

Problem 7 Direction of polarity

1. **F** is δ^- (3.98 against 2.20).
2. **O** is δ^- (3.44 against 2.55).
3. The arrows point from H to F and from C to O, toward the more electronegative atom.

Problem 8 Nitrate and resonance

1. N in the centre with one N=O and two N-O⁻. Formal charges: N = 5 - 0 - 4 = +1; double-bonded O = 0; each single-bonded O = -1. Total = +1 + 0 - 2 = -1.
2. **Three** resonance structures.
3. Bond order = (2 + 1 + 1)/3 = **1.33**.

Problem 9 Choosing the best structure

1. S-C≡N: S = 6 - 6 - 1 = -1; C = 0; N = 5 - 2 - 3 = 0. S=C=N: S = 0; C = 0; N = 5 - 4 - 2 = -1. S≡C-N: S = 6 - 2 - 3 = +1; C = 0; N = 5 - 6 - 1 = -2.
2. S≡C-N is least favourable: it has large formal charges (+1 and -2).
3. S=C=N puts -1 on N, which is more electronegative than S, so **S=C=N is slightly better**; both of the first two contribute.

Problem 10 Expanded octet in SF₆

1. Sulfur at the centre with six S-F single bonds; each F has three lone pairs.
2. **12 electrons** (six bonding pairs), an expanded octet allowed for period 3.
3. **Octahedral**, with **90°** angles.

Problem 11 Predict the shape

1. BF₃: trigonal planar electron geometry, trigonal planar shape, **sp²**.
2. SO₂: trigonal planar electron geometry, **bent** shape (one lone pair), **sp²**.
3. CH₄: tetrahedral electron geometry and shape, **sp³**.

Problem 12 Why do angles differ?

1. CH₄ has 0, NH₃ has 1 and H₂O has 2 lone pairs.
2. Lone pairs occupy more space than bonding pairs, so they push the bonding pairs together: the more lone pairs, the **smaller the angle**.

Problem 13 Is the molecule polar?

1. CO₂ is **nonpolar**: the two C=O dipoles cancel in the linear shape.
2. SO₂ is **polar**: the bent shape means the dipoles do not cancel.
3. CCl₄ is **nonpolar**: the four dipoles cancel by symmetry.
4. CHCl₃ is **polar**: H is different from Cl, so the dipoles do not cancel.

Problem 14 Sigma and pi bonds in acrylonitrile

1. C-H ×3, C=C one σ, C-C one σ, C≡N one σ gives **6 σ bonds**.
2. C=C has one π and C≡N has two π: **3 π bonds**.
3. a: sp²; b: sp²; c: **sp**.

Problem 15 Ozone

1. O=O⁺-O⁻ and O⁻-O⁺=O (central O bonded to both).
2. End O with a double bond: 6 - 4 - 2 = 0. Central O: 6 - 2 - 3 = +1. End O with a single bond: 6 - 6 - 1 = -1. Total 0.
3. Bond order = **1.5**; the angle is about **117°** (bent).

Problem 16 Xenon tetrafluoride

1. Four Xe-F bonds use 8 electrons; each F has 3 lone pairs, 4 × 6 = 24; total 32 so far. The remaining 36 - 32 = 4 electrons are **two lone pairs** on Xe.
2. Six domains: electron geometry **octahedral**; molecular shape **square planar**.
3. **sp³d²**. The four Xe-F dipoles cancel in a symmetric square, so the molecule is nonpolar.

Problem 17 Bond order in the nitrogen family

1. 945/160 = **5.9 times** stronger.
2. More shared pairs increase the electron density between the nuclei, increasing their attraction to the shared electrons and pulling them closer: 145 → 125 → 110 pm.
3. The triple bond is very strong (945 kJ/mol), so a large amount of energy must be supplied to break it. **That is why N₂ reacts only under extreme conditions.**

Problem 18 Dinitrogen monoxide

1. N≡N⁺-O⁻ and ⁻N=N⁺=O.
2. First: end N = 5 - 2 - 3 = 0; central N = 5 - 0 - 4 = +1; O = 6 - 6 - 1 = -1. Second: end N = 5 - 4 - 2 = -1; central N = +1; O = 6 - 4 - 2 = 0.
3. The first has the negative charge on O, the more electronegative atom, so it contributes **slightly more**. The molecule is **linear** (sp central N).

Problem 19 Chlorine trifluoride

1. Three Cl-F bonds use 6 electrons; the F atoms have 18; total 24. The remaining 4 electrons are **2 lone pairs** on Cl.
2. Five domains: electron geometry **trigonal bipyramidal**; molecular shape **T-shaped** (both lone pairs equatorial).
3. **sp³d**. The molecule is **polar** because the shape is asymmetric.

Problem 20 Nitrogen dioxide ions

1. NO_2^+ : $\text{O}=\text{N}^+=\text{O}$. NO_2^- : N with one double bond, one single bond and a lone pair.
2. NO_2^+ is **linear, 180°, sp** (two domains). NO_2^- is **bent, about 115°, sp²** (three domains).
3. NO_2^- has a lone pair on N, which repels the bonds and reduces the angle below 120°. NO_2^+ has no lone pair.