ANSWER KEY Answers and explanations appear after each question, marked by a bar at the left. Answers written in table cells are in *italics*.

ZornScience
zornscience.org

AP Chemistry Unit 2 Study Guide | ANSWER KEY

Compound Structure and Properties (Topics 2.1 to 2.7)

Show all work. When you explain, name the particles involved (ions, nuclei, electrons, bonding pairs, lone pairs) and describe how they interact.

Equations	Formal charge = (valence e^-) − (nonbonding e^-) − ½(bonding e^-)	Coulomb's law: $E \propto q_1 q_2 / r$	Bond order with resonance = total bonds ÷ number of bond positions
	2 electron domains: sp, 180°	3 electron domains: sp^2, 120°	4 electron domains: sp^3, 109.5°

Part 1. Types of Chemical Bonds (Topic 2.1)

1. Use the electronegativity (EN) values to classify each bond as **nonpolar covalent**, **polar covalent**, or **ionic**. For each polar or ionic bond, name the atom that carries the negative (or partial negative, δ−) charge.
EN values: H 2.20, C 2.55, N 3.04, O 3.44, F 3.98, Na 0.93, Mg 1.31, Cl 3.16

Bond	ΔEN	Bond type	Atom with negative charge
C–H	*0.35*	*nonpolar covalent (C and H share almost equally)*	*none (effectively nonpolar)*
O–H	*1.24*	*polar covalent*	*O*
N–Cl	*0.12*	*nonpolar covalent*	*none (very small difference)*
C–F	*1.43*	*polar covalent*	*F*
Na–Cl	*2.23*	*ionic (metal and nonmetal)*	*Cl, as Cl^-*
Mg–O	*2.13*	*ionic (metal and nonmetal)*	*O, as O^{2-}*

The larger the difference in EN, the more unequally the electrons are shared. The more electronegative atom pulls the shared electrons toward itself and gets the δ− charge.
Bonding is a continuum. Na–Cl and Mg–O are called ionic because they form between a metal and a nonmetal and their compounds have ionic properties, not just because of a ΔEN cutoff.

2. Rank the bonds H–F, H–Cl, H–Br, and H–I from **least** to **most** polar. Explain the trend using the structure of the halogen atoms.

H–I < H–Br < H–Cl < H–F
Electronegativity decreases down Group 17. Going down the group, the valence electrons are in higher shells, farther from the nucleus and more shielded by inner electrons, so the nucleus attracts a shared pair less strongly (Coulomb's law). F attracts the shared pair most, so H–F has the largest bond dipole.

3. Three solids were tested. Identify the type of bonding in each and cite the evidence that supports your claim.

Solid	Observations	Type of bonding	Evidence
X	White, brittle. Melts at 801 °C. Does not conduct as a solid. Conducts when melted or dissolved in water.	*ionic*	*Ions are locked in place in the solid but free to move when melted or dissolved. The high melting point comes from strong Coulombic attractions between ions.*
Y	Shiny and malleable. Melts at 1085 °C. Conducts as a solid and as a liquid.	*metallic*	*Delocalized valence electrons carry charge in any state, and layers of cations can slide without breaking the bonding.*
Z	Soft, waxy. Melts at 44 °C. Does not conduct in any state, even when melted.	*covalent (molecular solid)*	*Neutral molecules held by weak intermolecular forces: low melting point and no charged particles that can move.*

> The properties of a substance are the best evidence for its type of bonding.

4. Circle T (true) or F (false). If a statement is false, rewrite it so it is true.

(a) **T** **F** Any bond with an electronegativity difference greater than 1.7 is ionic, and every bond below that value is covalent.

> **False.** There is no sharp cutoff. Bonding is a continuum, and all polar bonds have some ionic character. Bonds between a metal and a nonmetal are generally ionic, and properties are the best evidence.

(b) **T** **F** In a polar covalent bond, the less electronegative atom develops a partial negative charge.

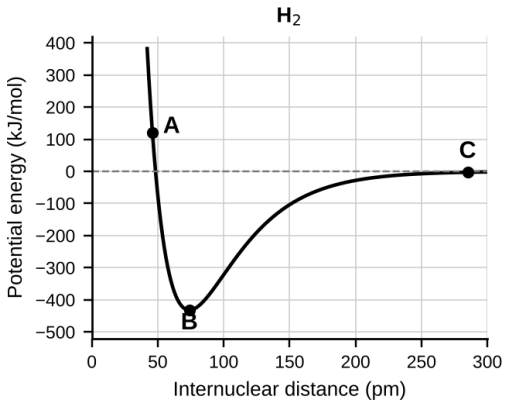
> **False.** The *more* electronegative atom pulls the shared electrons closer and develops the partial negative charge.

(c) **T** **F** In a metallic solid, the valence electrons are delocalized and not tied to any one atom.

> **True.** Metallic bonding is an array of metal cations in a "sea" of delocalized valence electrons.

Part 2. Intramolecular Force and Potential Energy (Topic 2.2)

5. The graph shows the potential energy of two hydrogen atoms as a function of the distance between their nuclei.



(a) What is the bond length of H_2? Explain how you know.

> About **74 pm**. The bond length is the distance where potential energy is lowest (point B).

(b) What is the bond energy of H_2?

> About **432 kJ/mol**. This is the energy needed to separate the atoms, from the bottom of the well (B) up to 0 kJ/mol (C).

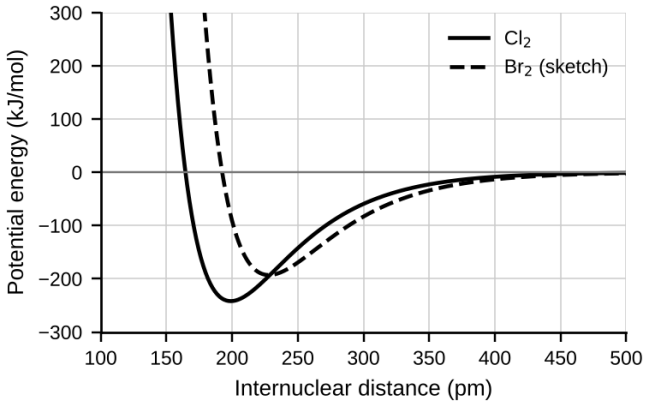
(c) Explain why the potential energy rises steeply at point A.

> At very short distances, the two positive nuclei (and the two electron clouds) repel each other strongly. Repulsion outweighs attraction, so potential energy increases sharply.

(d) Why is the potential energy close to zero at point C?

> The atoms are so far apart that they do not interact. There is no attraction or repulsion, so this is defined as zero potential energy.

6. The curve below is for Cl_2 (bond length 199 pm, bond energy 242 kJ/mol). On the same axes, sketch a curve for Br_2. Then justify the position and depth of your curve.



The Br_2 curve (dashed) should have its minimum farther to the right and should be shallower.
Br atoms are larger than Cl atoms (more occupied shells), so the bond is longer, about 228 pm. The shared electrons are farther from each nucleus and attracted less strongly, so the bond is weaker, about 193 kJ/mol.

7. N_2, O_2, and F_2 have bond lengths of 121 pm, 142 pm, and 110 pm and bond energies of 498 kJ/mol, 159 kJ/mol, and 945 kJ/mol (listed in no particular order). Complete the table by matching each molecule to its values.

Molecule	Bond order	Bond length (pm)	Bond energy (kJ/mol)
N_2	*3 (triple)*	*110*	*945*
O_2	*2 (double)*	*121*	*498*
F_2	*1 (single)*	*142*	*159*

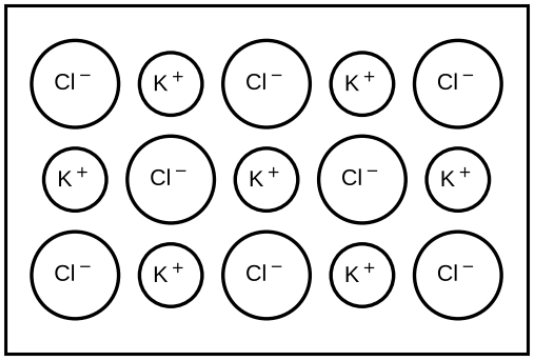
Higher bond order means more shared electron pairs between the nuclei, which pulls the atoms closer together (shorter bond) and takes more energy to break (larger bond energy).

8. Rank KCl, NaCl, CaO, and MgO from **weakest** to **strongest** attraction between their ions. Justify your ranking using Coulomb's law. (Ionic radii: K^+ 138 pm, Na^+ 102 pm, Ca^{2+} 100 pm, Mg^{2+} 72 pm, Cl^- 181 pm, O^{2-} 140 pm)

KCl < NaCl < CaO < MgO
Attraction is proportional to q_1q_2 and increases as r decreases. CaO and MgO have +2 and −2 ions, so q_1q_2 is 4 times larger than in KCl and NaCl. That is the biggest factor.
Among ions with the same charges, the smaller cation gives a smaller distance between ion centers: Na^+ is smaller than K^+, and Mg^{2+} is smaller than Ca^{2+}. So NaCl beats KCl and MgO beats CaO.

Part 3. Structure of Ionic Solids (Topic 2.3)

9. In the box, draw a particle diagram of solid potassium chloride, KCl, with at least 12 ions. Show the relative sizes of the ions (K^+ is smaller than Cl^-) and label each type of ion.



> Ions alternate in a regular, repeating pattern so each K^+ is surrounded by Cl^- ions and each Cl^- is surrounded by K^+ ions. This arrangement maximizes attractions between opposite charges and minimizes repulsions between like charges. The Cl^- circles should be drawn larger than the K^+ circles.

10. When an ionic crystal is struck with a hammer, it shatters instead of bending. Explain why at the particle level.

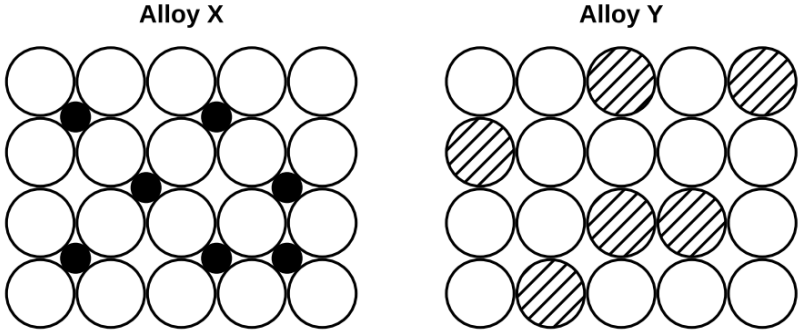
> The blow shifts one layer of ions relative to the next. After the shift, ions of the same charge line up next to each other. The strong repulsion between like charges pushes the layers apart, so the crystal breaks along that plane.

11. Complete the table. For each sample that conducts, name the particles that carry the charge.

Sample	Conducts electricity? (yes or no)	Particles that move to carry charge
NaCl(s)	*no*	*none: ions are held in fixed positions*
NaCl(l)	*yes*	*Na^+ and Cl^- ions*
NaCl(aq)	*yes*	*Na^+(aq) and Cl^-(aq) ions*
Cu(s)	*yes*	*delocalized valence electrons*
$C_{12}H_{22}O_{11}$(aq), sugar water	*no*	*none: neutral molecules, no ions*

Part 4. Structure of Metals and Alloys (Topic 2.4)

12. The diagrams model two alloys. (Atomic radii: Fe 126 pm, C 77 pm, Cu 128 pm, Zn 134 pm)



(a) Identify each alloy as **interstitial** or **substitutional**. Explain using the relative sizes of the atoms.

> **X is interstitial:** much smaller atoms fill the spaces (interstices) between the larger atoms. **Y is substitutional:** atoms of about the same size replace some of the host atoms in the lattice.

(b) Which diagram models steel (Fe with C), and which models brass (Cu with Zn)? Use the radii to justify.

> **Steel is X.** C (77 pm) is much smaller than Fe (126 pm), so C fits in the holes. **Brass is Y.** Zn (134 pm) is close in size to Cu (128 pm), so Zn takes the place of Cu atoms.

(c) Steel is harder and less malleable than pure iron. Explain why.

> The C atoms in the interstices interfere with the layers of Fe atoms, so the layers cannot slide past each other as easily. The lattice resists changing shape.

13. Use the electron sea model to explain why copper is (i) malleable and (ii) a good conductor of electricity.

> (i) Copper cations sit in a sea of delocalized valence electrons. When layers of cations slide past each other, the electron sea still surrounds them and holds them together, so the metal bends instead of breaking.
> (ii) The delocalized valence electrons are free to move through the whole solid, so they flow when a voltage is applied.

Part 5. Lewis Diagrams (Topic 2.5)

14. Complete the table. Draw a complete Lewis diagram for each species, showing every bond and lone pair.

Species	Total valence electrons	Lewis diagram	Lone pairs on central atom
PCl_3	*26*	*P in the center with single bonds to 3 Cl. One lone pair on P. Three lone pairs on each Cl.*	*1*
CO_2	*16*	*O=C=O. Two double bonds. Two lone pairs on each O.*	*0*
HCN	*10*	*H–C≡N. A single bond H–C and a triple bond C≡N. One lone pair on N.*	*0*
NH_4^+	*8*	*N in the center with single bonds to 4 H. No lone pairs. Draw brackets with a + charge.*	*0*
H_2CO (C is central)	*12*	*C in the center with single bonds to 2 H and a double bond to O. Two lone pairs on O.*	*0*

15. Draw the Lewis diagram for SF_4 (34 valence electrons). How many electrons surround the S atom? Explain why S can have this many electrons around it, but O cannot.

> S in the center with single bonds to 4 F atoms and **one lone pair** on S. Each F has 3 lone pairs.
> S is surrounded by **10 electrons** (4 bonding pairs and 1 lone pair), an expanded octet. Atoms in period 3 and beyond can hold more than 8 valence electrons. O is in period 2, where the valence shell (2s and 2p) can hold only 8 electrons.

Part 6. Resonance and Formal Charge (Topic 2.6)

16. The nitrate ion is NO_3^-.

(a) Draw all of the resonance structures for NO_3^- (24 valence electrons).

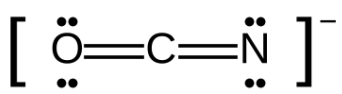
> Three equivalent structures. In each, N is central with one N=O double bond and two N–O single bonds. The double-bonded O has 2 lone pairs and each single-bonded O has 3 lone pairs. Connect the structures with double-headed arrows and put the whole set in brackets with a − charge.

(b) What is the N–O bond order? How does an N–O bond in nitrate compare in length to a typical N–O single bond and N=O double bond?

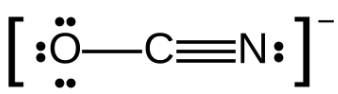
> Bond order = 4 bonds ÷ 3 positions = **4/3, or about 1.33**. All three N–O bonds are identical. Each is shorter than a single bond and longer than a double bond. No single Lewis structure shows this, which is why resonance is needed.

17. Two possible Lewis structures for the cyanate ion, OCN^-, are shown.

Structure I



Structure II



Atom	Formal charge in Structure I	Formal charge in Structure II
O	*6 − 4 − ½(4) = 0*	*6 − 6 − ½(2) = −1*
C	*4 − 0 − ½(8) = 0*	*4 − 0 − ½(8) = 0*
N	*5 − 4 − ½(4) = −1*	*5 − 2 − ½(6) = 0*

(a) Calculate the formal charge on each atom in both structures and record them in the table.

> See the table. In both structures the formal charges add up to −1, the charge on the ion.

(b) Which structure is the better model of OCN^-? Justify your choice using formal charge.

> **Structure II.** Both structures have formal charges as close to zero as possible (one atom at −1), but the best structure places the negative formal charge on the most electronegative atom. O is more electronegative than N, so the structure with −1 on O is preferred.

18. NO_2 has 17 valence electrons. Explain why no Lewis diagram can give every atom an octet, and describe one property of NO_2 that this helps explain.

> With an odd number of electrons, the electrons cannot all be paired, so one atom must have an unpaired electron and fewer than 8 electrons. This is a limitation of the Lewis model.
> Because of the unpaired electron, NO_2 is very reactive. Two NO_2 molecules join to form N_2O_4, which pairs the electrons ($2NO_2 \rightleftharpoons N_2O_4$).

Part 7. VSEPR and Hybridization (Topic 2.7)

19. Complete the table. Use "less than" when a lone pair compresses the ideal angle. For central atoms with more than 4 electron domains, hybridization is not needed.

Species	Electron domains on central atom	Molecular geometry	Bond angle(s)	Hybridization of central atom	Polar or nonpolar?
BF_3	*3*	*trigonal planar*	*120°*	*sp^2*	*nonpolar*
NH_3	*4*	*trigonal pyramidal*	*less than 109.5° (about 107°)*	*sp^3*	*polar*
H_2O	*4*	*bent*	*less than 109.5° (about 104.5°)*	*sp^3*	*polar*
SO_2	*3*	*bent*	*less than 120° (about 119°)*	*sp^2*	*polar*
PCl_5	*5*	*trigonal bipyramidal*	*90°, 120°, and 180°*	not needed	*nonpolar*
SF_4	*5*	*seesaw*	*less than 90° and less than 120°*	not needed	*polar*
XeF_4	*6*	*square planar*	*90° and 180°*	not needed	*nonpolar*

> A double bond counts as one electron domain (SO_2 has 2 bonding domains and 1 lone pair). Molecules are nonpolar when their shape is symmetric so the bond dipoles cancel (BF_3, PCl_5, XeF_4).

20. CH_4 has H–C–H bond angles of 109.5°, NH_3 has H–N–H angles of about 107°, and H_2O has an H–O–H angle of about 104.5°. All three central atoms have 4 electron domains. Explain the difference in bond angles.

> Lone pairs are held only by the central atom, so they spread out more and repel more strongly than bonding pairs. CH_4 has no lone pairs, NH_3 has 1, and H_2O has 2. Each added lone pair pushes the bonding pairs closer together, making the bond angle smaller.

21. Both CO_2 and SO_2 contain polar bonds. Explain why SO_2 is a polar molecule but CO_2 is not.

> CO_2 is linear (no lone pairs on C), so its two C=O bond dipoles point in opposite directions and cancel. SO_2 is bent because of the lone pair on S, so its bond dipoles do not cancel and the molecule has a net dipole moment.

22. Acrylonitrile has the structure $H_2C{=}CH{-}C{\equiv}N$. Number the carbon atoms C(1), C(2), and C(3) from left to right.

(a) How many sigma (σ) bonds and how many pi (π) bonds are in one molecule?

> **6 σ bonds and 3 π bonds.** Every bond has one σ bond (3 C–H, C=C, C–C, C≡N gives 6). The double bond adds 1 π bond and the triple bond adds 2 π bonds.

(b) Give the hybridization of C(1), C(2), and C(3).

> C(1): **sp^2** (3 domains). C(2): **sp^2** (3 domains). C(3): **sp** (2 domains).

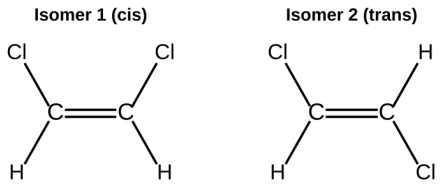
(c) Which is stronger, a σ bond or a π bond? Explain.

> A **σ bond**. Its orbitals overlap head-on, directly between the nuclei, which gives stronger overlap than the side-by-side overlap of a π bond.

23. Rank the carbon-oxygen bonds in CO, CO_2, and CO_3^{2-} from **shortest** to **longest**. Justify using bond order.

> **$CO < CO_2 < CO_3^{2-}$**
> Bond orders: CO is 3 (triple), CO_2 is 2 (double), and CO_3^{2-} is 4/3 because of resonance among three equivalent structures. Higher bond order gives a shorter, stronger bond.

24. The two isomers of 1,2-dichloroethene are shown.



(a) Explain why the isomers do not turn into each other at room temperature.

> The π bond in C=C prevents rotation around the bond. Turning one isomer into the other would require breaking the π bond.

(b) Which isomer is polar? Explain.

> **Isomer 1 (cis).** Both C–Cl dipoles point toward the same side, so they do not cancel. In the trans isomer, the dipoles point in opposite directions and cancel.