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Topic 3.1: Intermolecular and Interparticle Forces

Worksheet · AP Chemistry

UNIT 3 · PROPERTIES OF SUBSTANCES AND MIXTURES

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How to Use This Worksheet

Work through the sections in order. Section A checks the ideas, Section B builds fluency with electron counts, data and rankings, and Section C mixes everything the way an exam would.

Use the same three steps for every question. First, decide from the structure whether each particle is polar or nonpolar. Second, list every intermolecular force acting between the particles, not just the strongest. Third, rank the total attraction using the strength drivers: electron count and polarizability, surface contact, hydrogen bond donors and acceptors, ionic charge and size.

Do not look at the answer key until you have attempted every question. The key gives full working, so compare your reasoning, not just your final answer.

Section A: Recall and Concepts

A1. Dry ice, solid CO_2, sublimes at -78 °C, turning straight into a gas. (a) Name the type of intermolecular force that is overcome when dry ice sublimes. (b) State what happens to the C=O covalent bonds inside the CO_2 molecules during sublimation. (c) Explain why sublimation counts as a physical change even though the solid disappears completely.

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A2. (a) Name the four intermolecular and interparticle forces studied in this topic, and state the kinds of particles each one acts between. (b) State which of the four is present in every substance, and explain in one sentence where it comes from.

A3. For each of the following substances, state whether it can form hydrogen bonds with identical molecules in its pure liquid: NH_3, H_2O, CH_3F, H_2S, CH_4. For each substance that cannot, state exactly which requirement it fails.

A4. (a) Explain why a sample of liquid HCl experiences both London dispersion forces and dipole-dipole interactions, rather than dipole-dipole interactions alone. (b) State the two features a molecule must have before any dipole-dipole interactions can occur between its molecules.

A5. Water molecules cluster around dissolved ions. (a) Name the force acting between a Na^+ ion and the water molecules surrounding it, and state which end of each water molecule points toward the ion. (b) Mg^{2+} has an ionic radius of 72 pm and Na^+ an ionic radius of 102 pm. State which ion attracts water dipoles more strongly, and give both reasons.

Section B: Core Practice

B1. The noble gases are monatomic and nonpolar, with boiling points: He -269 °C, Ne -246 °C, Ar -186 °C, Kr -153 °C, Xe -108 °C. (a) Name the only intermolecular force present in liquid argon. (b) Calculate the number of electrons in an atom of argon and in an atom of krypton. (c) Explain, in terms of polarizability, why krypton boils 33 °C higher than argon. (d) Two unidentified noble gases

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boil at -186 °C and -152 °C. Identify both elements, and state which one has the more polarizable electron cloud.

B2. The diatomic halogens boil at: F_2 -188 °C, Cl_2 -34 °C, Br_2 59 °C, I_2 184 °C. (a) Calculate the total number of electrons in a molecule of F_2 and in a molecule of I_2. (b) Explain the trend in boiling points from F_2 to I_2. (c) A student explains the trend by writing "the covalent bonds get stronger down the group." Use the bond enthalpies Cl-Cl 242 kJ/mol and I-I 151 kJ/mol to identify the two errors in that sentence.

B3. Hydrogen sulfide, H_2S, boils at -60 °C. Silane, SiH_4, boils at -112 °C. (a) Calculate the total number of electrons in each molecule, and the molar mass of each compound. (b) State the shape of each molecule and whether it is polar or nonpolar. (c) List every intermolecular force present in each pure liquid. (d) Explain why H_2S boils 52 °C higher than SiH_4 even though the two compounds have the same electron count and similar molar masses.

B4. Three isomers of C_5H_{12} boil at: pentane 36 °C, 2-methylbutane 28 °C, neopentane 10 °C. ΔHvap is 25.8 kJ/mol for pentane and 22.7 kJ/mol for neopentane. (a) Calculate the total number of electrons in a C_5H_{12} molecule, and state why neither molar mass nor electron count can explain the differences between these boiling points. (b) Rank the three isomers by the surface area their molecules present to neighbouring molecules, and connect the ranking to the boiling point order. (c) Calculate the difference between the two enthalpies of vaporization, and the percentage by which neopentane's value is lower than pentane's. (d) State what that difference measures, in terms of intermolecular forces.

B5. Ethanol, CH_3CH_2OH, and dimethyl ether, CH_3OCH_3, share the molecular formula C_2H_6O and the molar mass 46.07 g/mol. Ethanol boils at 78 °C and dimethyl ether at -24 °C. (a) List every intermolecular force present in each pure liquid. (b) Explain why dimethyl ether cannot form

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hydrogen bonds with identical molecules even though its oxygen atom carries lone pairs. (c) Explain the 102 °C gap between the two boiling points.

B6. Three +1 cations sit in water: Li^+ with radius 76 pm, Na^+ with radius 102 pm, and K^+ with radius 138 pm. (a) Rank the three ions by the strength of their ion-dipole attraction to water, strongest first, and justify the ranking. (b) A simple model treats ion-dipole strength as proportional to charge ÷ radius. Calculate charge ÷ radius for Li^+, Na^+, K^+ and for Mg^{2+} (charge +2, radius 72 pm), and state which ion the model ranks strongest overall. (c) Explain why the model places Mg^{2+} above Li^+ even though the two ions are nearly the same size.

Section C: Exam Style

C1. The hydrogen halides have these boiling points: HF 20 °C, HCl -85 °C, HBr -67 °C, HI -35 °C. (a) Calculate the number of electrons in each of the four molecules. (b) Explain why the boiling points rise from HCl to HBr to HI. (c) HF has the fewest electrons of the four, yet it boils 105 °C above HCl. Explain this anomaly fully, naming the force responsible and the structural feature that makes it possible. (d) State which of the four liquids exhibit hydrogen bonding, and which exhibit dipole-dipole interactions.

C2. At room temperature, Cl_2 is a gas, Br_2 is a liquid and I_2 is a solid. All three are nonpolar diatomic molecules. Hydrogen chloride, HCl, is polar and is a gas at room temperature, boiling at -85 °C. (a) State the intermolecular forces present in Cl_2, Br_2 and I_2. (b) The electron counts are 34 for Cl_2, 70 for Br_2 and 106 for I_2. Use them to explain the three different states at room temperature. (c) A student writes: "Dipole-dipole forces are always stronger than London dispersion forces, so a polar substance always boils higher than a nonpolar one." Use the substances in this question to show that the statement is false, and state the condition under which dispersion forces overtake dipole-dipole forces.

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C3. Three small hydrides, each containing exactly 10 electrons, have these data: CH_4 boils at -161 °C, ΔHvap 8.2 kJ/mol; NH_3 boils at -33 °C, ΔHvap 23.4 kJ/mol; H_2O boils at 100 °C, ΔHvap 40.7 kJ/mol. (a) List every intermolecular force present in each of the three liquids. (b) Calculate the ratio of the enthalpy of vaporization of water to that of methane, and state what the size of the ratio tells you. (c) Both NH_3 and H_2O form hydrogen bonds, yet water boils 133 °C higher than ammonia. Explain this in terms of the hydrogen bond donors and lone pair acceptors each molecule carries. (d) A student writes: "Water molecules are held together only by hydrogen bonds, because hydrogen bonding is so much stronger than dispersion forces." Correct the statement.

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Answer Key

A1. (a) London dispersion forces. CO_2 is linear and nonpolar, so dispersion forces are the only attractions between its molecules. (b) Nothing: the C=O bonds stay intact, and every CO_2 molecule that leaves the solid is complete and unchanged. (c) A physical change separates particles without altering them: only the spacing of the molecules changes as the forces between them are overcome. No bond within a molecule breaks and no new substance forms, so the change is physical, not chemical.

A2. (a) London dispersion forces, between all atoms and molecules. Dipole-dipole interactions, between polar molecules. Hydrogen bonding, between molecules carrying a hydrogen bonded to N, O or F and a lone pair on N, O or F. Ion-dipole forces, between an ion and a polar molecule. (b) London dispersion forces are present in every substance: any electron cloud is uneven at any instant, creating a temporary dipole that induces a matching dipole in a neighbour.

A3. NH_3 can: it has N-H bonds to donate with and a lone pair on nitrogen to accept with. H_2O can: it has O-H bonds and two lone pairs on oxygen. CH_3F cannot: it fails the donor requirement, because its hydrogens are bonded to carbon, not fluorine, so none is sufficiently δ+, even though its fluorine carries lone pairs. H_2S cannot: it fails the donor requirement because sulfur is not one of N, O or F, so its hydrogen is not δ+ enough. CH_4 cannot: it fails both requirements, having no N, O or F atom and no lone pair.

A4. (a) Dispersion forces come from electron motion, which every molecule has, so they act between HCl molecules as between any others. HCl is also polar, so dipole-dipole attractions act in addition. Polarity adds a term; it does not replace one. (b) At least one polar bond, and a geometry in which the bond dipoles do not cancel, so the molecule carries a net permanent dipole.

A5. (a) Ion-dipole forces. The oxygen end of each water molecule points toward Na^+, because oxygen carries the δ- end of the dipole and opposite charges attract. (b) Mg^{2+}, for two reinforcing reasons. Its charge is +2 rather than +1, so each dipole is pulled by twice the charge. And its radius is smaller, 72 pm against 102 pm, so its charge is more concentrated and water can approach closer, where attraction is stronger.

B1. (a) London dispersion forces only: argon atoms are monatomic and nonpolar, so no other force on the list can act. (b) Argon is element 18: 18 electrons per atom. Krypton is element 36: 36 electrons per atom. (c) Krypton's cloud is larger, holds twice the electrons, and is held less tightly, so it distorts more easily: krypton is more polarizable. Its temporary dipoles and dispersion attractions are stronger, so more thermal energy is needed to separate its atoms into a gas. Hence -153 °C rather than -186 °C, a difference of -153 - (-186) = 33 °C. (d) The -186 °C gas is argon; the -152 °C gas is krypton, whose measured boiling point is -153 °C. Krypton has the more polarizable cloud.

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B2. (a) F_2: 2 × 9 = 18 electrons. I_2: 2 × 53 = 106 electrons. (b) All four molecules are nonpolar, so only dispersion forces act. Down the group the electron count climbs and each cloud grows larger and more polarizable, so the dispersion attractions strengthen steadily and the boiling point climbs from -188 °C to 184 °C. (c) First error: boiling separates molecules from each other, so the trend is set by forces between molecules, and the covalent bonds inside them are never broken. Second error: the facts run the other way. Bond enthalpies fall from Cl-Cl at 242 kJ/mol to I-I at 151 kJ/mol, so the bonds weaken down the group while the boiling point rises by 184 - (-34) = 218 °C.

B3. (a) H_2S: 16 + 2 = 18 electrons. SiH_4: 14 + 4 = 18 electrons. Molar masses: H_2S = 2 × 1.008 + 32.06 = 34.08 g/mol; SiH_4 = 28.09 + 4 × 1.008 = 32.12 g/mol. (b) H_2S is bent, like water, and its bond dipoles do not cancel: polar. SiH_4 is tetrahedral and symmetrical, so its dipoles cancel: nonpolar. (c) Liquid H_2S: dispersion forces and dipole-dipole interactions; it does not hydrogen bond, because sulfur is not N, O or F. Liquid SiH_4: dispersion forces only. (d) The two tie on the dispersion drivers: identical electron counts and molar masses differing by only (34.08 - 32.12) ÷ 32.12 × 100 = 6.1%, so their dispersion forces are closely matched. H_2S then adds dipole-dipole attractions on top of that shared base, because it is polar and SiH_4 is not, and its boiling point sits -60 - (-112) = 52 °C higher.

B4. (a) A C_5H_{12} molecule contains 5 × 6 + 12 × 1 = 42 electrons. All three isomers share the same formula, so their molar masses (72.15 g/mol) and electron counts are identical by construction. Quantities that are exactly equal cannot produce three different boiling points; the cause must be something that differs: molecular shape. (b) Pentane, the unbranched chain, presents the largest contact surface; then 2-methylbutane, with one branch; then neopentane, compact and near spherical, with the least. The boiling points follow the same order, 36 °C, 28 °C, 10 °C: greater contact means stronger total dispersion attraction. (c) Difference: 25.8 - 22.7 = 3.1 kJ/mol. Percentage: 3.1 ÷ 25.8 × 100 = 12.0% lower for neopentane. (d) ΔHvap is the energy needed to pull one mole of molecules completely apart, so the 3.1 kJ/mol difference measures how much less strongly a mole of neopentane molecules holds together than a mole of pentane molecules: the energy price of the contact area that branching removes.

B5. (a) Ethanol: dispersion forces, dipole-dipole interactions and hydrogen bonding. Dimethyl ether: dispersion forces and dipole-dipole interactions only. (b) A hydrogen bond needs a donor: a hydrogen bonded directly to N, O or F. In dimethyl ether every hydrogen is bonded to carbon, so although its oxygen lone pairs make each molecule a potential acceptor, the pure liquid contains no donors at all, and no hydrogen bonds form. (c) The two liquids are matched on size: same formula, same molar mass of 46.07 g/mol, same electron count of 26, so their dispersion forces are similar, and both, being polar, add dipole-dipole attractions. Ethanol adds a third term dimethyl ether cannot build: a hydrogen bonding network through its O-H group, worth 78 - (-24) = 102 °C of boiling point.

B6. (a) Li^+ strongest, then Na^+, then K^+. All carry +1, so size decides: the smaller the ion, the more concentrated its charge and the closer a water dipole can approach, so attraction strengthens as

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radius falls from 138 pm to 76 pm. (b) Li^+: 1 ÷ 76 = 0.0132 per pm. Na^+: 1 ÷ 102 = 0.0098 per pm. K^+: 1 ÷ 138 = 0.0072 per pm. Mg^{2+}: 2 ÷ 72 = 0.0278 per pm. The model ranks Mg^{2+} strongest overall, at 0.0278 ÷ 0.0098 = 2.8 times the value for Na^+. (c) The radii are nearly equal, 72 pm against 76 pm, so the size term ties. The charge term does not: +2 against +1 doubles the charge ÷ radius value, 0.0278 against 0.0132, a factor of about 2.1. When size ties, charge decides.

C1. (a) HF: 1 + 9 = 10 electrons. HCl: 1 + 17 = 18. HBr: 1 + 35 = 36. HI: 1 + 53 = 54. (b) All three heavier molecules are polar, so each has dispersion forces plus dipole-dipole interactions. What changes is the electron count, climbing 18, 36, 54: each cloud is larger and more polarizable, so the dispersion component grows steadily, lifting the boiling point from -85 °C to -67 °C to -35 °C. (c) HF hydrogen bonds. Its hydrogen is bonded directly to fluorine, the most electronegative element, making it exceptionally δ+, and fluorine carries lone pairs to accept with: both halves of the requirement are met in their strongest form. Hydrogen bonding far outweighs HF's dispersion deficit from having only 10 electrons, lifting its boiling point 20 - (-85) = 105 °C above HCl's. (d) Hydrogen bonding: HF only. Dipole-dipole interactions: all four, since all four molecules are polar; in HF, hydrogen bonding is the dominant special case.

C2. (a) London dispersion forces only, in all three: each molecule pairs two identical atoms, so it is nonpolar, with no permanent dipole and no hydrogen bonded to N, O or F. (b) Dispersion is the whole story, and its strength tracks electron count. With 34 electrons, Cl_2's attractions are too weak to hold a liquid at room temperature: a gas. With 70, Br_2's hold a liquid, but not a solid. With 106, I_2's lock its molecules into a solid. (c) The counter example is in the question: nonpolar I_2, held together by dispersion forces alone, is a solid at room temperature, while polar HCl is a gas until -85 °C. Dispersion forces overtake dipole-dipole forces when the nonpolar particle's electron cloud is large and polarizable enough, that is, when its electron count is high, because dispersion strength grows with cloud size while dipole-dipole strength does not.

C3. (a) CH_4: dispersion forces only; it is nonpolar and contains no N, O or F. NH_3: dispersion forces, dipole-dipole interactions and hydrogen bonding. H_2O: dispersion forces, dipole-dipole interactions and hydrogen bonding. (b) Ratio: 40.7 ÷ 8.2 = 4.96, about 5.0. Vaporising a mole of water takes about five times the energy of vaporising a mole of methane with the same electron count and so similar dispersion forces. The ratio prices water's extra forces, above all hydrogen bonding. (c) Count both halves of hydrogen bonding. Water carries two O-H donors and two lone pairs, so each molecule can join up to four hydrogen bonds, building a fully connected three dimensional network. Ammonia carries three N-H donors but only one lone pair, and every hydrogen bond consumes one of each, so the lone pair shortage caps how connected each ammonia molecule can be. Its network is sparser, its attraction weaker, and its boiling point lower by 100 - (-33) = 133 °C. (d) The correction: water is held together by dispersion forces and dipole-dipole interactions as well as hydrogen bonds, since dispersion forces act between all molecules without exception. Hydrogen bonding is the strongest of the three terms, not the only one.

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