

London, dipole and hydrogen bond attractions between molecules

Chemistry grades 11-12 (A Level · IB · Class 11-12) · Worksheet answers

Each answer names the forces, compares the number of electrons first and then adds dipole or hydrogen bonding effects. In explanations, always say that intermolecular forces are overcome, and that covalent bonds are not broken on boiling.

Level: Chemistry grades 11-12 (A Level · IB · Class 11-12) · **Curriculum:** IB S2.2

Problem 1 Bonds or forces?

1. Only the **intermolecular forces (hydrogen bonds)** between water molecules weaken as ice melts.
2. The covalent bonds **stay intact**; the molecules separate but remain H₂O.
3. Intermolecular forces are 10 to 100 times weaker than covalent bonds, so far less energy is needed to overcome them.

Problem 2 Names of the forces

1. **London forces.**
2. **Permanent dipole-dipole** forces (plus London forces).
3. **Hydrogen bonding.**
4. **London forces** (non-polar molecule).

Problem 3 Origin of London forces

1. The order is **B, C, A.**
2. An **instantaneous (temporary) dipole.**

Problem 4 Counting electrons

1. CH₄: $6 + 4 = \mathbf{10}$.
2. C₂H₆: $12 + 6 = \mathbf{18}$.
3. H₂O: $2 + 8 = \mathbf{10}$.
4. Cl₂: $2 \times 17 = \mathbf{34}$.

Problem 5 Who can form hydrogen bonds?

1. **NH₃, CH₃OH and HF** can form hydrogen bonds.
2. CH₃OCH₃ has no hydrogen atom bonded to N, O or F (the H atoms are all bonded to carbon).

Problem 6 Kelvin and Celsius

1. He: $4 - 273 = -269\text{ °C}$; Ne: $27 - 273 = -246\text{ °C}$; Xe: $165 - 273 = -108\text{ °C}$.
2. **Xenon**, since it has the most electrons (54), the largest electron cloud and the strongest London forces.

Problem 7 Boiling points of alkanes

1. The boiling point **rises** with the number of carbon atoms, with smaller increases for longer chains.
2. A longer molecule has more electrons and more surface contact, so London forces are stronger and more energy is needed to separate molecules.
3. The last increase was 33 °C , so continuing the trend gives **about 100 °C** for heptane.

Problem 8 Branched isomers

1. Both have **London forces** only (non-polar alkanes, same electrons).
2. The branched molecule is more compact and has a smaller surface area, so there is less contact between molecules, the London forces are weaker, and **less energy is needed to boil it**.

Problem 9 The halogens at room temperature

1. Cl_2 : $2 \times 17 = 34$; Br_2 : $2 \times 35 = 70$; I_2 : $2 \times 53 = 106$.
2. As the electrons increase, London forces strengthen, so the molecules are held together more strongly. At 25 °C only Cl_2 has weak enough forces to be a gas, Br_2 is a liquid and I_2 is a solid.

Problem 10 Fluorine and hydrogen chloride

1. Both molecules have the **same number of electrons (18)**, so the London forces are about equal.
2. HCl is polar and has permanent dipole-dipole forces in addition, while F_2 is non-polar. So **HCl needs more energy** to boil.

Problem 11 Propanone and butane

1. Propanone has **permanent dipole-dipole** forces (the $\text{C}=\text{O}$ bond is polar).
2. The permanent dipoles add to the London forces and outweigh the small difference in electrons (32 against 34).
3. It has no hydrogen atom bonded to oxygen, nitrogen or fluorine; all H atoms are bonded to carbon.

Problem 12 Ethanol and methoxymethane

1. Ethanol: $\text{CH}_3\text{CH}_2\text{OH}$; methoxymethane: CH_3OCH_3 .
2. Ethanol: **hydrogen bonding**; methoxymethane: **dipole-dipole** forces.
3. Both have 26 electrons, so the London forces are similar. Hydrogen bonds are stronger than dipole-dipole forces, so ethanol needs more energy to boil.

Problem 13 Hydrogen halides

1. From HCl to HI the electrons increase (18, 36, 54), so London forces strengthen and the boiling point rises.
2. HF has H bonded to F, so it forms **hydrogen bonds**, which make it boil far higher than the trend predicts.

Problem 14 A hydrogen-bonded chain

1. Both A and B are **hydrogen bonds**.
2. On A: H of the left molecule is $\delta+$ and F of the right molecule is $\delta-$.
3. The strongest bond has an angle close to **180°** .

Problem 15 Glycerol and propan-1-ol

1. Glycerol has **3** O-H groups; propan-1-ol has **1**.
2. Each glycerol molecule can form many hydrogen bonds with neighbours, so the molecules are tangled and cannot slide past one another easily. The liquid is more viscous.

Problem 16 Hydrogen fluoride, water and ammonia

1. HF: 1 hydrogen and 3 lone pairs; H_2O : 2 hydrogens and 2 lone pairs; NH_3 : 3 hydrogens and 1 lone pair.
2. Water has two hydrogens and two lone pairs, so each molecule can form up to **four hydrogen bonds** (two as donor, two as acceptor) in a 3D network. This needs the most energy to break.
3. In NH_3 nitrogen is less electronegative than oxygen or fluorine, so the N-H bond is less polar and the hydrogen bonds are weaker. It has only one lone pair, so it forms fewer bonds, which gives the lowest boiling point.

Problem 17 Energy to boil water

1. $n = 250 \div 18.0 = 13.9 \text{ mol}$; energy = $13.9 \times 40.7 = \mathbf{565 \text{ kJ}}$.
2. One mole of H_2O has 2 mol of O-H bonds: $2 \times 463 = \mathbf{926 \text{ kJ}}$.
3. Boiling needs only 40.7 kJ per mole, about 4 % of 926 kJ. So the **covalent O-H bonds stay intact**; only intermolecular hydrogen bonds are broken.

Problem 18 Mixing ethanol, water and hexane

1. Water: **hydrogen bonding**; hexane: **London forces**.
2. Ethanol has an O-H group, so it forms hydrogen bonds with water of similar strength to those it breaks, so the two liquids mix.
3. Hexane cannot form hydrogen bonds with water. The strong water-water hydrogen bonds would be replaced by weak attractions, which costs energy, so the liquids separate.

Problem 19 Ranking three liquids

1. Butane **-1 °C**, propanone **56 °C**, propan-1-ol **97 °C**.
2. Butane has only London forces. Propanone adds permanent dipole-dipole forces. Propan-1-ol has an O-H group and forms hydrogen bonds, the strongest of the three.

Problem 20 Groups 14 and 15

1. Group 14 hydrides are non-polar tetrahedral molecules, so only London forces act. The boiling point rises from -162 °C to -52 °C as the electrons increase.
2. **NH₃ (-33 °C)** is out of line: it can form hydrogen bonds, whereas PH₃ cannot.
3. Carbon, silicon, germanium and tin are not electronegative enough, so their hydrides have no N-H, O-H or F-H bond and no hydrogen bonding.