Solutions

Each answer shows the steps, then the **result**.

2.1

- nitrogen, oxygen (and fluorine)

2.2

- the London dispersion / instantaneous dipole–induced dipole force

2.3

- The two $C = O$ dipoles are equal in size and point in opposite directions; the molecule is linear, so they cancel and there is no overall dipole

2.4

- A hydrogen atom bonded to N, O or F; and a lone pair on an N, O or F atom of another molecule

2.5

- 2-methylpropane is branched, so its molecules have less surface contact with each other; the van der Waals' forces between them are therefore weaker and less energy is needed to separate them

3.1 B

- Hydrogen bonds hold ice in an open lattice, making it less dense than water

3.2 C

- Water has hydrogen bonding, the strongest of the four

3.3

(a)

- water molecules are held by hydrogen bonds; much energy is needed to break these, so the boiling point is high

(b)

- down the group atoms have more electrons, so stronger London forces, needing more energy to overcome — higher boiling points

3.4

- a dashed hydrogen bond from a $\delta+$ H on one molecule to a lone pair on the $\delta-$ O of the other; both molecules drawn with $\delta+$ on H and $\delta-$ on O; a clear lone pair shown on the accepting oxygen

3.5

(a)

- van der Waals' forces; permanent dipole–dipole forces

(b)

- NH_3 has hydrogen bonding, because H is bonded to the small, highly electronegative nitrogen, which also carries a lone pair; PH_3 has only van der Waals' and dipole–dipole forces, since phosphorus is not electronegative enough; hydrogen bonds are much stronger, so far more energy is needed to separate ammonia molecules, outweighing PH_3's greater number of electrons

(c)

- Oxygen is more electronegative than nitrogen and each water molecule has two hydrogens and two lone pairs, so water forms more hydrogen bonds per molecule, and stronger ones

(d)

- The hydrogen bonds and other intermolecular forces between the molecules are broken; the covalent O–H bonds within the molecules are NOT broken

(e)

- In ice each molecule is held by hydrogen bonds to four others in a rigid tetrahedral arrangement; this open lattice contains gaps, so the same number of molecules occupies a larger volume; the density is therefore lower than that of liquid water and ice floats