Answers to questions

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Question 1. Water (H_2O) has a melting point of 0^oC and a boiling point of 100^oC :
(a) Compare the figures to the melting points of two other group VI dihydrides, H_2S and H_2Se.

H_2S has melting point of -82^oC and boiling point -60^oC
H_2Se has melting point of -66^oC and boiling point -41^oC (figures taken from wikipedia)

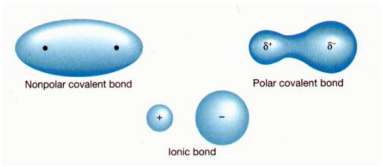
b. Analyse the differences in melting and boiling points of the above substances in relation to the relative electronegativity of sulphur, selenium and oxygen.

Electronegativity, measured using the Pauling scale, gives us the following table.

Element(X)	Electronegativity	Boiling point of H_2X (^oC)
S	2.58	-82
Se	2.55	-66
O	3.44	0
H	2.2	-

So no clear pattern - there are other things involved. However the steep jump in electronegativity between Oxygen vs S & Se could suggest that the large difference in melting points is partly because the hydrogen bonds are not as strong in H_2S or H_2Se.

Recall the reason water has such a high melting and boiling point for a small covalent molecule (compare to things like methane which boils at -161 ^oC or the diatomic oxygen we breath which boils at -183 ^oC). Water has polar covalent bonding, so it has a stronger attraction to neighbouring molecules than with just pure Van der Waals.



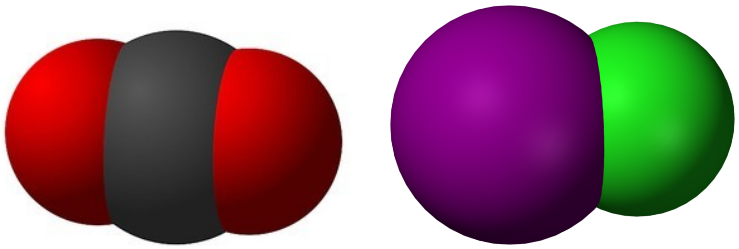
The difference in electronegativity between the atoms in a bond will determine how polar, or how non-polar the bond between them is.

To illustrate this let us compare two diatomic molecular compounds: Bromine monochloride (BrCl) and Iodine monochloride (ICl).

Chlorine has electronegativity of 3.16, Bromine of 2.96, and Iodine of 2.66. So the difference in electronegativity between Cl - Br is 0.2 and between Cl - I is 0.5. BrCl is fairly non-polar and has melting point of -66 ^oC; in contrast ICl has a much higher polarity and thus can form stronger polar bonds giving it a melting point between 13 and 27 ^oC depending on the structure.

So given that the difference in electronegativity can determine the degree of polarity in a **bond**, then we can guess that Oxygen, with a higher electronegativity than S or Se, will make **more polar** bonds with hydrogen. So the hydrogen bonding in solid water will be 'stronger' than the hydrogen bonding in either H_2S or H_2Se.

Note: while the difference in electronegativity between two bonding atoms can predict the degree of polarity **in that bond**, it cannot tell you the degree of polarity **in the molecule**. For example the electronegativity difference between Carbon and Oxygen is actually more than in ICl, however CO^2 is SYMMETRICAL - so even though the carbon-oxygen bonds are polar, overall the molecule is not polar.



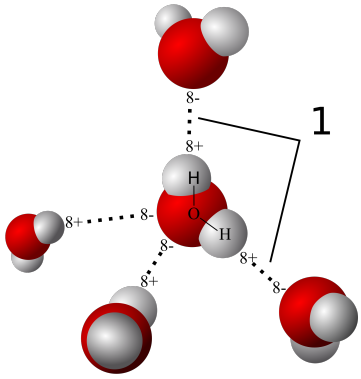
Practice question: can the difference in electronegativity between Hydrogen and Fluorine HF help you predict whether or not the molecule exhibits polar bonding? What about the difference in electronegativity between Lithium and Fluorine?

`https://www.youtube.com/watch?v=126N4hox9YA`

0.1 Question 2

Water (H2O) is less dense as a solid than a liquid. Analyse how the forces acting between water molecules change as water solidifies.

So recall water has the ability to form **hydrogen bonds**. These hydrogen bonds are a very strong kind of covalent bonding, much stronger than Van der Waals. However because the bond is polar, the water molecules prefer to be in certain orientations.



This means when the water gets cold, it likes to form a **lattice** structure. At normal pressure this lattice structure has a specific size, determined by the most favourable distances between the slightly positive Hydrogens and the slightly negative Oxygens of neighbouring molecules. It just happens that this lattice is **less dense** than liquid water.

When water is melted (a liquid) the molecules are rotating and swimming about, the lattice structure is gone. So the molecules get much closer.

Useful links:
`https://www.youtube.com/watch?v=6s0b_keOiOU`
`https://www.youtube.com/watch?v=UE1-iBJja4Y`
`http://www1.lsbu.ac.uk/water/water_hydrogen_bonding.html` `https://www.khanacademy.org/science/biology/water-acids-and-bases/hydrogen-bonding-in-water/a/hydrogen-bonding-in-water`

0.2 Question 3

Diamond and graphite are both made of pure carbon. Diamond is the hardest known substance and cannot conduct electricity, while graphite is very soft and can conduct electricity. Analyse how the intermolecular forces acting within these substances lead to the different properties

Simply put, diamond has a different structure to graphite. In diamond there are covalent bonds between each carbon atom and the four carbon atoms neighbouring it. These covalent bonds are known as 'sigma' bonds, and the valence electrons involved in each bond are 'bounded', i.e. they will remain floating about between the two carbon atoms. This mega-structure, usually referred to as a 'covalent network', is incredibly strong. Why? Because covalent bonds are

strong, and distorting the structure would involve breaking some covalent bonds - which takes a lot of energy.
So diamond's hardness is explained because distorting the structure requires a lot of energy to break a lot of strong C-C covalent bonds. But also diamond's non-conductivity makes sense because there are no free-moving charge carriers. This property is similar to some other covalent network structures, e.g. Silicon Oxide, which is also hard and a non-conductor. http://www.avogadro.co.uk/structure/chemstruc/network/g-molecular.htm

In contrast, graphite, has a very different structure. It has sheets of carbon atoms. The intra-sheet bonding is strong, but the sheets are only held together by weak van der waal forces, so the sheets can slide past each other, bend, and flex. This is why graphite is so soft. But also electrons can flow along the sheets, meaning graphite can conduct electricity.

http://www.chemguide.co.uk/atoms/structures/giantcov.html
http://www.bbc.co.uk/schools/gcsebitesize/science/add_ocr_gateway/chemical_economics/nanochemistryrev1.shtml
https://www.khanacademy.org/partner-content/crash-course1/crash-course-chemistry/v/chem34-network-solids-carbon

0.3 Question 4

Analyse the influence of molecular size on the intermolecular forces acting between molecules of benzene and naphthalene. You should explain how these forces affect the melting points of these substances.

The only intermolecular forces **between** benzene molecules are London Dispersion Forces (Van der Waals), the same as naphthalene. London Dispersion forces are in general much weaker than polar covalent bonds. However the magnitude of LDF's scales with the size of the electron cloud surrounding the molecule. Bigger cloud - bigger wobble - bigger temporary dipoles - greater London Dispersion Forces.

Naphtalene is a bigger molecule than Benzene, in fact it's got 4 extra carbons and 2 extra hydrogens, it is nearly twice the size of Benzene. This means it has a much bigger electron cloud to 'wobble' and create the London dispersion attraction.

Bigger cloud $\rightarrow$ bigger wobble $\rightarrow$ larger temporary dipoles $\rightarrow$ greater attraction for the London Dispersion Forces $\rightarrow$ more energy required to overcome those attractions $\rightarrow$ higher melting and boiling point. Sure enough Naphthalene melts at 80 and boils at 218 ^{o}C whereas Benzene melts at 5 and boils at 80 ^{o}C.

https://www.youtube.com/watch?v=G1jGeeSWhXY
http://www.chemguide.co.uk/atoms/bonding/vdw.html