

## Chapter 13 Questions

1. a) Gas molecules can experience Van der Waal forces. These forces are too weak to hold particles together at room temperature and atmospheric pressure. By reducing the temperature (reduced molecular motion) and increasing the pressure (forces molecules closer together) Van der Waal forces are able to pull gas molecules together and cause condensation to a liquid.

b) An ideal gas has negligible attractive forces. Gases that behave in an ideal manner have a very low Van der Waal force of attraction and are therefore hard to condense.

2. ~~Water~~


2. Bonds in methane: covalent bond, very strong bond (all covalent bonds are), shared electrons in a molecular orbital

Bonds between methane molecules: Van der Waal forces, weak, slight interaction between electrons and nuclei of adjacent molecules

3. a) Krypton: Krypton has  $36 e^-$ , while Ne has only  $10 e^-$ . The greater the number of electrons, the greater the Van der Waal attractions between molecules.

b) The larger the molecule, the greater the Van der Waal forces. Larger molecules have more electrons (increases Van der Waal forces) and greater surface areas (increases Van der Waal forces). The increase in Van der Waal forces increases the boiling point and decreases the vapour pressure (molecules have more difficulty escaping the force of attraction to become a "free" gas molecule).

4. ~~6~~  $6 H$   hexane

~~4~~  $4$   2,2-dimethylbutane

H has a greater surface area contact between molecules, therefore increased V.d.W. forces, therefore increased boiling point

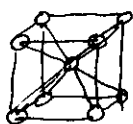
5. Diamond is a 3-dimensional network solid with extremely strong ~~&~~ covalent bonds. Diamond powder is very hard and can contain sharp edges, ~~graphite~~ therefore very abrasive. Graphite is a 2-dimensional network solid. Layers of flat covalently bonded hexagonal networks of carbon atoms are held together by weak Van der Waal forces. It is easy for the planes of carbon atoms to slip easily past one another. This gives graphite its lubricating property.
6.  $\text{SnCl}_4 < \text{SnBr}_4 < \text{SnI}_4$  As the number of electrons in the outer atoms increases, so does the the Van der Waal force of increases there by raising the melting point.
7. Graphite breaks down to a black powder. Tale is white in colour which is less noticeable ~~of~~ on clothing or skin. Tale is also softer and smoother feeling than graphite.

8. a) Metals are good conductors of electricity (easy to test) and are easy to work (malleable and ductile) (also easy to test)
- b) Metals have "delocalized" or "free moving" electrons which ~~can~~ allow good conductivity (electrons move freely through the metal) The workability of metals can be explained by the structure of a solid metal. Positive charged metal ions held together in an electron soup allows for metal ions to slip past one another and yet maintain a strong metallic bond.
9. In a plane in graphite electrons are able to travel through a hexagonal network of conjugated ~~to~~ double bonds. This results in delocalized electrons that can conduct an electrical current (similar to a metal). Electrical conduction perpendicular to a plane is very poor in graphite because electrons are not able to travel freely between planes (metals conduct equally in all directions)

10a) Atoms with lower ionization energy are able to give up electrons more easily thus increasing the amount of free moving electrons available to cause metallic bonding. In addition atoms with empty orbitals in which free moving electrons can travel make good metals.

b) Sc will have a higher heat of vapourization than K because it has stronger metallic bonds. Sc can lose  $3e^-$  per atom leaving behind  $Sc^{3+}$  ions. K can only lose  $1e^-$  per atom leaving behind  $K^+$  ions. The greater number of free moving electrons and greater ionic charge in Sc results in a stronger metallic bond.

11a)



8 nearest neighbours

c) see page 344 - 12 nearest neighbours

e) The crystal structure is related to the properties of the metal, particularly density

12. The energy released by the hydration of  $Na^+$  and  $Cl^-$  ions provides sufficient energy to overcome the lattice energy of the ionic solid

13. Water has an unusually high boiling point due to hydrogen bonding. Hydrogen bonding does not occur in  $H_2S$ ,  $H_2Se$  or  $H_2Te$  (the electronegativity of S, Se and Te is insufficient to expose the hydrogen proton and allow hydrogen bonding). The boiling point of  $H_2S < H_2Se < H_2Te$  due to increasing Van der Waal forces in the larger atoms.  $H_2O$  does not agree with this trend due to stronger hydrogen bonding (hydrogen bonds have moderate strength) - see pg. 348

14.  $PH_3$  will have a lower boiling point than  $NH_3$ .  $NH_3$  will form hydrogen bonds (moderate strength bonds) while  $PH_3$  will only have Van der Waal forces (some dipole interactions as well) which is significantly weaker than hydrogen bonding.

15. Water - the universal solvent - is central in biological systems. The solvent properties of water are largely due to hydrogen bonding.  
 $H_2O$  / Sugar       $H_2O$  / protein

16. a) Molecular Solid: discrete molecules held together by Van der Waal forces or hydrogen bonds  
eg. water, candle wax,
- b) Network Solid: three dimensional - interconnected network of covalently bonded <sup>atoms</sup> resulting in very hard crystal structures  
eg. diamond, quartz (silicon oxide)
- c) Network Solid: two dimensional - planes of interconnected covalently bonded atoms, planes held together by weak Van der Waal forces  
eg. graphite, talc
- d) Metals: positive metal atoms held together by a "sea" of electrons, metal atoms are packed together in a crystal structure  
eg. sodium, iron, gold, etc.
- e) Ionic Solid: strong electrostatic forces of attraction between oppositely charged ions creates a regular arrangement of ions in a crystal lattice of positive and negative ions  
eg. NaCl, LiF, MgO, KCl

17. Molecular Solids: low melting and boiling points, soft compounds, may be crystalline or irregular,  
Network Solid: high melting points and boiling points, hard compounds, crystalline structures  
Network Solids (two dimensional): soft brittle compounds, high melting and boiling points, electrical conduction parallel to planes.  
Metals: high melting and boiling points, good electrical conductivity, malleable, ductile  
Ionic Solids: high melting and boiling points, hard, crystalline compounds, soluble in polar solvents ( $H_2O$ )
18. Conductivity in solid silver is due to free moving electrons within the metallic solid. Electrons are small and flow easily. Conductivity in molten AgCl is due to free moving  $Ag^+$  particles.  $Ag^+$  ions are larger particles and do not flow as easily as electrons in a metal, therefore less conduction. In solid AgCl there are no free moving charge particles (free moving charged particles are necessary for electrical conduction) therefore no conduction occurs in solid AgCl

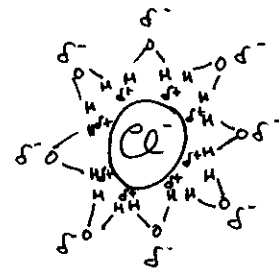
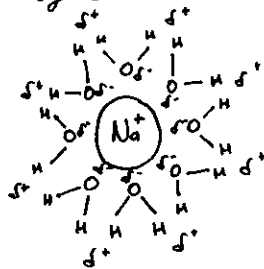
19. Ionic crystals will "snap" along planar orientations of particles in the crystal lattice. Ionic bonds cannot be bent, therefore ionic crystals break when stressed. Metallic ~~solids~~ bonds are able to bend somewhat. The metallic bonds allows metal ions to slip past one another without interrupting the bond. Therefore metals will deform when stressed.

20. ~~Lat He(l) V.D.W. B) Fe(s) metallic c) diamond s) network solid~~

20.	Vander Waals	Covalent Network Solid	Metallic	Ionic
a) He(l)		c) diamond(s)	b) Fe(s)	e) KCl
j) C <sub>2</sub> H <sub>2</sub> (l)		g) Si (s)	f) Ag(s)	m) MgO
k) SO <sub>2</sub> (l)		h) SiO <sub>2</sub> (s)		
		l) SiC(s)		

d) CH<sub>4</sub>(g) and I<sub>2</sub>(g) contain covalent ~~net~~ bonds within the molecule but do not have any forces acting between molecules

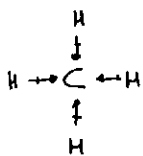
21. NaCl dissolves in water through a process of hydration. Water molecules (a polar molecule) orientate themselves around positive Na<sup>+</sup> and negative Cl<sup>-</sup> ions in such a way as to maximize + and - attractions. Each ion becomes surrounded by a sheath of water molecules during the hydration process



22. H<sub>2</sub>SO<sub>4</sub> dissociates in water to form H<sup>+</sup>, H<sup>+</sup> and SO<sub>4</sub><sup>2-</sup> ions. Each of these ions is hydrated by water molecules. The resulting hydration liberates a considerable amount of heat.

23. Water is able to hydrate  $\text{Na}^+$  and  $\text{Cl}^-$  ions. (polar molecule)  
 The energy released is more than the energy stored in the crystal lattice structure of  $\text{NaCl}$  (s) baseline and Benzene are ~~not~~ <sup>non</sup> polar compounds and therefore do not interact strongly with  $\text{Na}^+$  and  $\text{Cl}^-$  ions. Because there is no strong interaction between solvent and solute (like there is between water and  $\text{NaCl}$ ) the lattice of  $\text{NaCl}$  cannot be broken apart.

24.



$\text{C-H}$  bonds are slightly polarized but all four polarizations cancel out due to the tetrahedral shape of the molecule (non-polar molecule)



$\text{C-F}$  bonds are so fairly polar. The shape of the molecule results in a net polarization of the ~~CH3~~  $\text{H}_3\text{CF}$  molecule. This molecule can have a dipole dipole interaction with other ~~CH3~~  $\text{H}_3\text{CF}$  molecules. Therefore  $\text{H}_3\text{CF}$  has a greater attraction between

molecules than  $\text{CH}_4$  (Van der Waals forces only) and therefore has a higher boiling point

25.  $\text{O}=\overset{\ominus}{\text{C}}=\overset{\oplus}{\text{O}}$  small molecule with no net polarization.  $\text{CO}_2$  solidifies at  $-78^\circ\text{C}$  since  $\text{CO}_2$  molecules have only slight attraction between molecules

$(\text{SiO}_2)_n$  Each silicon atom forms single covalent bonds with four ~~silicon~~ oxygen atoms which in turn covalently bond to another silicon atom.

The result is a network solid (quartz) not unlike diamond

26. Since silicon does not easily form double bonds it is unlikely that silicon could form a two dimensional covalent network with conjugated double bonds. A silicon analogue of graphite is unlikely.