

The interactions between the units in an aggregate are what lead to the physical and chemical characteristics of the substance.

Aggregates are usually crystalline solids.

Substances are usually classified in terms of the unit bonding that makes up the aggregate.

1 – Ionic Substances –
structural units are oppositely charged ions and the bond is ionic.

2 – Molecular Substances – (both polar and non-polar)

eg]non-polar – polar –

structural units are and the attractive forces between the molecules are called intermolecular forces.

3 – Network or Macromolecular solids – gigantic networks of atoms held together by covalent bonds.

4 – Metallic Substances – eg]

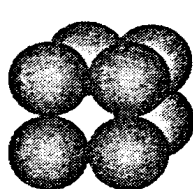
The structural unit is the atomic kernel (positive ion). These are held together by metallic bonds. (sea of electrons)

1 - Ionic Substances - Is it proper to talk about an NaCl molecule?

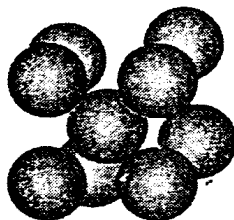
Properties of Ionic Substances

- A - MP and BP, volatility
- B - electricity when solid
- C - electricity when liquid
- D - in water to form (electrolytic)solutions
- E - brittle

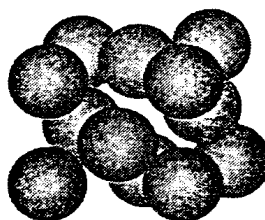
- the above properties can almost all be attributed to the high strength of the ionic bond.
- the solid phase of ionic substances are crystalline (repetitive unit cells). The unit cell is simply the most basic repeating unit. The position of the atom or ion in the unit cell is called a lattice point and the 3-D array of lattice points is called the lattice structure.
- unit cells - simple, face centred cubic and body centred cubic(determined by X-ray diffraction)



Simple cubic



Body-centred cubic



Face-centred cubic

The crystal lattice energy is

- the crystal lattice energy of ionic substances

- the high crystal lattice energy leads to high MP, BP and low volatility – (see property A)

Are the melting and boiling points of all ionic substances really high?

This difference in MP must be because the crystal lattice energies of different ionic substances are quite dissimilar, but why are they dissimilar?

Physics tells us that the force of attraction between two static charges is given as:

so for larger ionic radii –

so for larger charges –

Why do solid ionic crystals not conduct? (property B)

Why do melted ionic substances act as moderate conductors? (property C)

Why do solvated ionic substances conduct? (property D)

Why are ionic crystals brittle? (property E) (Hint: think of the packing)

2 – Molecular Crystals.

In general the intramolecular forces (those within the molecule) are much greater than the intermolecular forces (those between molecules) in the molecular crystal.

The lattice points are occupied by neutral molecules. (polar or non-polar)

There is a much wider variation in the properties of molecular substance than ionic substance because

Examples of molecular solids include:

In general:

A – neither the solid crystals nor the liquids conduct electricity

B – These molecular crystals have very low BP, MP and high volatility

C – the solids tend to be soft and waxy

D – there is a large amount of energy required after melting or boiling to decompose the substances chemically.

Why don't molecular substances conduct when solid or liquid? (property A)

Why do molecular solids have low BP, MP and high volatility? (property B)

Why are molecular solids soft and waxy? (property C)

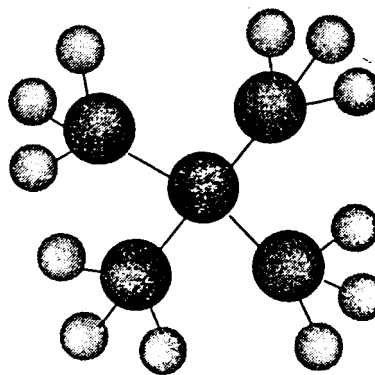
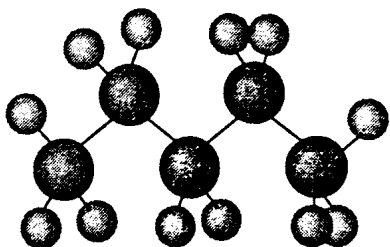
Why does it take so much energy after melting or boiling to decompose the substance? (property D)

How does shape of a molecule affect the properties?

normal pentane
 C_5H_{12}

vs.

neo pentane
 C_5H_{12}



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2b – What are these so called weak intermolecular forces anyway?

Dipole Forces

Van der Waals Forces

- these occur between every type of molecule (even 0 electronegativity difference molecules)
- they are the result of temporarily induced dipoles due to the fluctuation in electron charge density

- covalent bonds are 100's of times stronger than Van der Waals
- the magnitude of the Van der Waals forces therefore depends on

Molecule	Number of Electrons	Boiling Point (°C)
CH ₄	10	
SiH ₄	18	
GeH ₄	36	
SnH ₄	54	

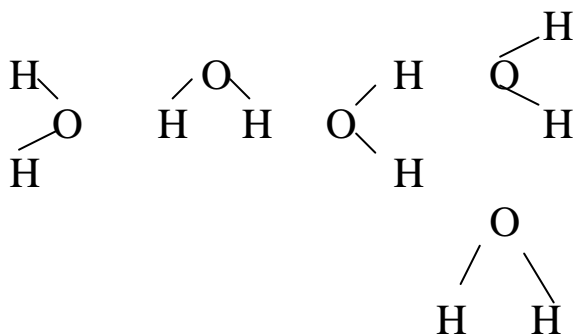
Hydrogen Bonds - Examine the following:

EN diff	Form	#e	BP	EN diff	Form	#e	BP	EN diff	Form	#e	BP
2.19	PH ₃	18	-85	2.58	H ₂ S	18	-60	3.10	HCl	18	-85
2.18	AsH ₃	36	-55	2.55	H ₂ Se	36	-42	2.96	HBr	36	-67
2.05	SbH ₃	54	-17	2.1	H ₂ Te	54	-2	2.66	HI	54	-36

The higher boiling point of the first row indicates

Hydrogen Bond – the relatively strong intermolecular force that results from electrostatic interaction between a hydrogen of one molecule and a highly electronegative atom of another molecule.

The high electronegativity results



Hydrogen bonds are approximately 10 times as strong as Van der Waals forces

Can be used to explain why water is more dense than ice (thankfully) and to explain the high surface tension of water and biomolecular shapes.

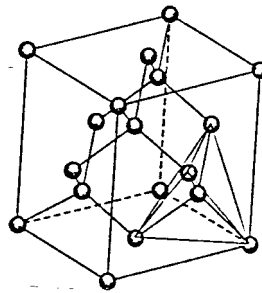
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3 - Network or Macromolecular Solids

- one, two or three dimensional networks of atoms

3D Network Solids

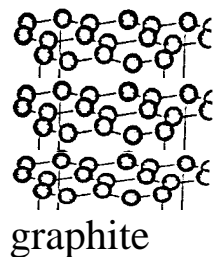
- eg] diamond, quartz, silicon carbide (carborundum)
 - High MP (ionics have a high MP but networks have very high MP)
 - Extreme hardness
 - In diamond all C-C bond angles are 109.28° and all C-C bonds are equal strength
 - A diamond is one very large molecule extended three dimensionally



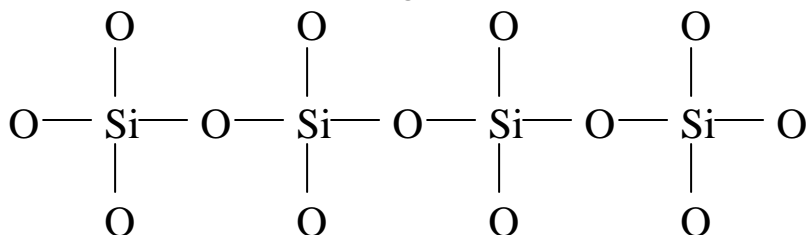
- The network of atoms (each C bonded to four others) accounts for the extreme hardness and high BP.
- When the highly directional bonds are broken by an applied stress, they cannot be reformed and the crystal cleaves.

2D Network Solids

- more than one type of bond exists
- graphite (pure carbon) – the bonds between the plane can be broken easily so the planes tend to slide over each other making graphite a good lubricant
- graphite conducts in one direction – along the π bond plane not perpendicular to it
- soft due to weak bonds between



1D Network Solids – eg] asbestos



4 - Metallic Crystals or Metallic Bonds

- the properties of metals cannot be accounted for by ionic or covalent bonds

Metals in general have the following properties:

- Lustre (reflectivity)
- Good conductors (heat and electricity)
- Workability (malleability, ductility)
- electron emission by heat or light
- covalent bonds are one directional and their fixed length would resist the workability (therefore metals are not bonded covalently)

- ionic bonds are extremely resistant to deformation (therefore metals are not bonded ionically)

- Van der Waals forces are much too weak to explain the stability of metals

- We must therefore have a different type of bond called a metallic bond - positive atomic kernels in a sea of free moving electrons.

Lustre - the free electrons oscillate at the same frequency as incoming light and because they then become accelerating charges they emit light of the same incident frequency making it look reflective.

Conductivity - the tiny free electrons can move readily under applied voltage. At high temperatures the atomic kernel vibration interferes with electron flow.

Heat Conductivity - again the free electrons can carry off excess kinetic energy from one area of the metal to another.

Workability - the non-directional bond nature allows one plane of atomic kernels to slide over another plane with no change to the immediate environment of the atomic kernel. This results in no added internal stress, so the metal does not crack.

Electron Emission - when enough heat energy is added, an electron may gain enough energy to be bounced out of the lattice. When light of the proper frequency interacts with an electron, the energy is sufficient to bump the electron out.