

Molecular Architecture - How to Use this Package

This package contains a tremendous amount of information about all attractive forces between atoms. This is basically the backbone of knowledge that leads to why matter appears the way it does. In other words, all physical properties such as melting point, boiling point, crystallinity, hardness, malleability, ductility, solubility, conductivity plus others.

Page #1: classifies all matter:

Molecule Type: macro vs discrete covalent

Macromolecules: occur whenever ionic or metallic bonds are present, or whenever covalent bonds link "endlessly" such as in diamond (a covalent network solid)

Discrete Covalent Molecules: occur whenever a small and limited number of atoms are linked together by covalent bonds

Intramolecular Force: occurs within a molecule

Intermolecular Force: occurs between discrete covalent molecules only

Page #2: describes in detail the three different types of intramolecular forces. Keep in mind that ionic and metallic bonds always result in macromolecules. Covalent bonds can form either macromolecules or discrete covalent molecules.

Page #3 and #4: describe in detail the physical properties of ionic, metallic and covalent network solids. This information does not apply to any small discrete covalent molecules such as the molecules studied in organic chemistry

Page #5: describes in detail both the attractive force and some physical properties for the three intermolecular forces that exist between discrete covalent molecules such as the molecules studied in organic chemistry

Page #6: summary of chapter #4 from the text

Page #7: a more complete classification scheme to include some exceptions

Read all pages through very carefully.

Additional reading: chapter 13 from the old text

Additional questions: # 1 to 31 starting on page 351 from old text

Classification of Matter According to Forces

		All Matter					
Molecule Type		Macromolecules		Discrete Covalent Molecules			
Intramolecular Force Type		Covalent Bond	Ionic Bond	Metallic Bond	Covalent Bond Only		
Intermolecular Force Type	N.A.	N.A.	N.A.	N.A.	Hydrogen Bond	Dipole interaction	van der Waals Force
Relative Strength	all about the same - 100						
Examples of solids	diamond (C _n), Quartz (SiO ₂)	table salt (NaCl), MgS, K ₂ O, AlCl ₃	gold, silver, iron, alloys such as steel or titanium- aluminum alloy		ice, benzoic acid (a carboxylic acid)	?	wax (an alkane), moth balls (an alkene)
Examples of liquids	very rare at room temperature (mercury is an example of metal), solids of ionic and metallic solids go through a liquid phase when heated, during which aspect of the bond are still in place solid of covalent solids frequently turn directly to a vapour (at very high temperatures in excess of 3000 °C)				water, ethyl alcohol, acetic acid	ether, formaldehyde, acetone, carbon disulphide	hexane, 1-hexene, cyclohexene, benzene (an aromatic ring), gasoline (mixed hydrocarbons)
Examples of gases	once a gas, all bonding forces have been overcome				once a gas, all intermolecular forces have been overcome (no longer present), however, the intramolecular covalent bonds are still as present as ever, i.e water vapour is still H ₂ O in units		

Three Bond Types! (Intramolecular Forces)

	Covalent Bonds	Ionic Bonds	Metallic Bonds
smallest units	neutral atoms (may be partially charged)	positive and negative ions	positive metal ions and free moving electrons
nature of bonding force	mutual electrostatic force of attraction between bonding pairs of electrons and the nuclear charge of both atoms involved in the covalent bond quantum mechanical waveforms for electrons dictate what will be a stable bonding arrangement (i.e. octet rule)	electrostatic force of attraction between positively and negatively charged ions	mutual electrostatic force of attraction between free moving electrons and metallic ions, electrons wander through empty valence shells
required elements	a non-metal element above the stairs (include hydrogen) bonded with itself or another element above the stairs (include hydrogen) between atoms that have medium to high electronegativity and ionization energy (i.e above the stairs and hydrogen)	one element with low ionization energy which loses an electron (or electrons) to an element of high electronegativity generally occurs between a metal element (low ionization energy) and a non-metal element (high electronegativity) can also involve covalently bonded polyatomic ions (i.e. NH_4^+ , SO_4^{2-})	one (or more in the case of alloys) element with low ionization energy and empty valence shells low ionization energy means the element can lose electrons easily which then travel freely throughout the empty valence shell orbitals throughout the metal
behaviour of electrons involved in bonding	locked into covalent bonds and cannot move "localized" ("delocalization" can occur in 2-d network solids such as graphite or organic molecules containing pi bonds) differences in electronegativity between bonding elements can slightly shift bonding electrons <u>within</u> the bond creating polarized covalent bonds	locked in place around individual ions and cannot move "localized"	free to wander throughout the empty valence shell orbitals of the metal ions, can easily wander from one ion to the next "delocalized"
crystal structure	geometric arrangement (called a crystal lattice) of neutral atoms linked by covalent bonds geometry is based on covalent bonding principles	geometric arrangement (called a crystal lattice) of positive and negative ions geometry is based on a lowest energy arrangement in which (net attractions)-(net repulsion) is a maximum	geometric arrangement of metal ions packed together to minimize volume packing arrangements: -body centered cubic: 8 neighbours -hexagonal close packed 12 -cubic close packed 12
strength of bond considerations	single, double and triple covalent bonds	magnitude of ionic charges, arrangement of ions in crystal	magnitude of metallic ion charge and number of electrons contributed to "electron soup", size of ions

Physical Properties of Ionic, Covalent Network and Metallic Solids

	Ionic Solids (Salts)	Covalent Network Solids	Metallic Solids
melting and boiling point	very high to high melting and boiling points: due to strong intramolecular forces of attraction (electrostatic forces of attraction between positive and negative ions) that require a lot of thermal energy to overcome	very high to high melting and boiling points: due to strong intramolecular forces of attraction (covalent bonds) that require a lot of thermal energy to overcome (do not confuse with discrete molecular compounds which have low to moderate melting and boiling points)	very high to moderate: the metallic bond strength can vary due to metal ion charge, ionic radius and number of electrons donated to the "electron soup" that holds the metal ions together, stronger metallic bonds of course require more thermal energy to overcome
crystal structure	highly structured due to lattice arrangement, lattice arrangement is a result of minimum energy considerations that maximize attractions between ions of opposite charge while minimizing repulsions between ions of like charge, anomalies are rare	highly structured due to lattice arrangement, lattice arrangement is a result of covalent bonding geometry producing three dimensional network giving a very strong crystal structure with few anomalies	moderate degree of structure usually not visible to the eye, the metallic bond does not require a specific geometry, crystal packing arrangements (i.e. body centered cubic, hexagonal close packed and cubic close packed) are merely way of packing ions together, crystal packing is prone to anomalies
response to physical stress or strain	may shatter: once the crystal lattice is disturbed, the ionic bonds will break if the ions are dislocated may cleave: the crystal lattice could fracture along a plane of ions, thus breaking ionic bonds along a flat face	may shatter: once the crystal lattice is disturbed, the covalent bond break may cleave: the crystal lattice could fracture along a plane of atoms, thus breaking covalent bonds along a flat face	malleable and ductile: will bend and stretch, metal ions can slide past each other without interrupting the metallic bond (the electrons move very fast and freely throughout the metal)



continued

solubility	<p>variable solubility in water and other <u>polar solvents</u>: electrostatic interaction between regions of partial charge on solvent molecules and the charge on ions can be sufficient to overcome the lattice energy of the ionic crystal, in water the dissolving process is called solvation or hydration, lattice energy > hydration energy insoluble ionic compound hydration energy > lattice energy soluble ionic compound</p>	<p>insoluble: neutral atoms in the crystal have a high latticed energy, there are no solvents that can interact successful with neutral atoms to overcome the high lattice energy, lattice energy must be overcome for dissolving to occur (lattice energy is associated with the force of attraction that holds a crystal together in a crystal lattice)</p>	<p>insoluble: free moving electrons are not soluble in any normal solvent (i.e. electrons cannot be hydrated), therefore metals cannot dissolve even though metal ions could be hydrated many metals are soluble in mercury metal: mercury metal can accommodate free moving electrons thus allowing metals to dissolve</p>
electrical conductivity	<p>solid crystals are non-conductive: electrons and ions are locked in place, no free moving charged particles, therefore no conduction molten (liquid) state are conductive: positive and negative ions are able to move, free moving charged particles allow for conduction (note: electrons are still locked into the ions) ionic solutions are conductive: free moving positive and negative ions in solution are able to carry a current</p>	<p>three dimensional network solids are non-conductive: electrons are localized in covalent bonds, no free moving charged particles, therefore no conduction two dimensional network solids (graphite) are conductive within a plane but none conductive between planes: electronic resonance within a plane allows for the movement of electrons through a plane, free moving charged particles, therefore conductive</p>	<p>conductive in solid and liquid state: free moving electrons traveling in the empty valence orbitals of the metal ions provide excellent conductivity</p>
molecular size	<p>macromolecule: unspecified large size</p>	<p>macromolecule: unspecified large size</p>	<p>macromolecule: unspecified large size</p>
examples	<p>NaCl, LiF,</p>	<p>diamond (C_n), quartz (SiO₄), LiN, Si, SiC</p>	<p>gold, silver, iron, steel alloys</p>

Physical Properties of Molecular Solids – Based on Intermolecular Forces

	Van Der Waals Forces	Dipole Interactions	Hydrogen Bond
relative strength	1 kcal/mol \approx 4 kJ/mol (can add up to more in larger molecules)	3–5 kcal/mol \approx 13–21 kJ/mol (per interaction)	10 kcal/mol \approx 42 kJ/mol (per interaction)
nature of force	electrostatic force of attraction between electrons in one molecule with nuclear charges in a neighboring molecules	electrostatic force of attraction between regions of partial positive or negative charge on one molecule with regions of partial negative or positive charge respectively on a neighboring molecules	electrostatic force of attraction between the positive charge of an "exposed" proton on one molecule and the negative charge of a lone pair on a neighboring molecule
required circumstance	occurs between any discrete covalent molecules but is often overshadowed by a stronger intermolecular force do not consider V. D. W. forces unless no other forces are operating	requires discrete covalent molecules with elements that differ in electronegativity in order to generate regions of partial charge requires a molecular geometry that does not cancel out regions of partial charge (i.e. unsymmetrical molecules)	require a hydrogen that is bonded to the very electronegative elements N, O, or F in order to create the exposed proton <u>AND</u> interaction with the lone pair on another N, O, or F in order to create the stronger interaction of a true hydrogen bond only N, O, or F are electronegative enough to "expose" a proton and only N, O, or F have a dense enough lone pair orbital to interact with the exposed proton to form a "bond"
conditions that alter strength	total number of electrons in molecule: the greater the number of electrons in the molecule the greater the v.d.w. force molecular shape: good surface contact between molecules increases force particularly in liquids, good packing geometry increases force in a solid	magnitude of partial charges: the greater the difference in electronegativity, the greater the partial charge, the greater the force of attraction geometry: a geometry that maximizes the availability of partial charge will enhance the force of attraction	all hydrogen bonds are roughly equivalent in strength, see above for required conditions
crystal structure of solid	nil to slight: usually fairly mushy like a wax	slight to considerable: depends packing geometry in the crystal	moderate to highly crystallized: depends on packing geometry in the crystal, ice forms highly crystalline structure under the right conditions
melting and boiling points	low for the size of the molecule	low to moderate	can be fairly high for the size of the molecule
examples	alkanes, alkenes, alkynes, aromatic rings	ketones, aldehydes, ethers	carboxylic acids, alcohols, water, sucrose, glucose, fructose

Chapter #4 Suggested Reading and Selected Questions

Section 4.1: be able to draw Lewis dot diagrams using techniques learned in class. Some tips:

- halogens always bond with simple single covalent bonds
- if extra electrons are present (i.e. a negative ion), the extra electrons will always be present on the most electronegative element (usually oxygen, which will now have seven electrons and behave like a halogen)
- use a balance of single coordinate covalent* bonds and double bonds to satisfy the octet rule if possible
- use expanded valence shells if necessary (this occurs frequently with halogens)
- use valence shells with less than an octet when necessary

* a single coordinate covalent bond is one in which one atom donates both electrons

Section 4.2: be familiar with:

- sp^3 , sp^2 and sp hybridization
- σ and π bonding and relation to single, double and triple covalent bonds
- figures 9 through 16

Section 4.3: VSEPR Theory will be covered thoroughly in class.

Read this section to make sure that you can understand the language being used.

Section 4.4: Read carefully and make notes as necessary, including the blue shaded areas.

Section 4.5: Read carefully and make notes as necessary, including the blue shaded areas. This section contains information about the three types of intermolecular forces. Please note that the London force is often referred to as a van der Waals force.

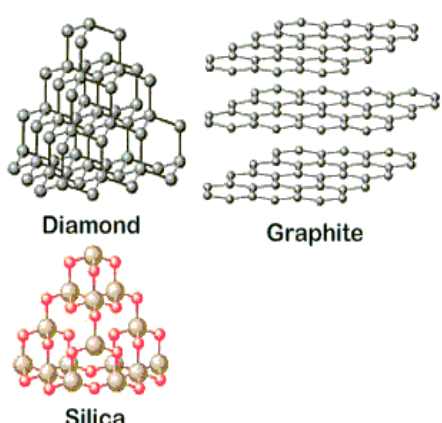
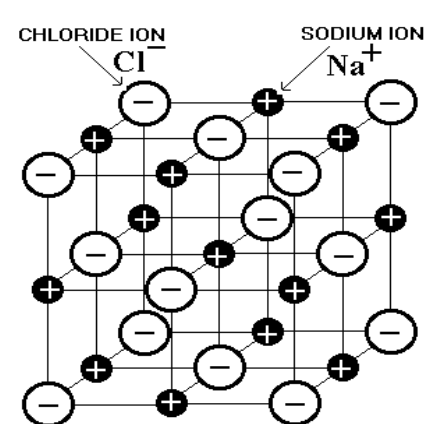
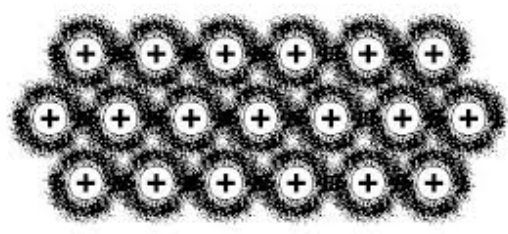
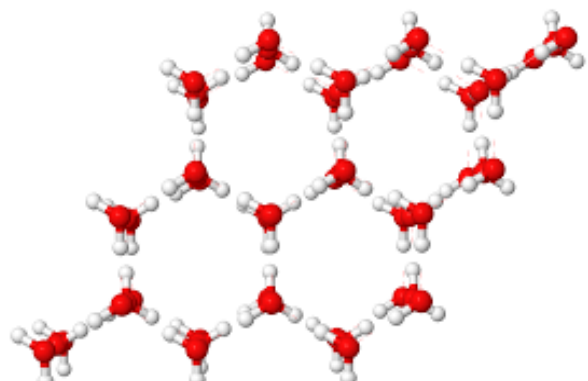
Section 4.6: Make careful notes, including explanations of physical properties. Note that ionic crystals, metallic crystals and covalent network crystals are macromolecules and have only intramolecular forces at work. Molecular crystals are crystals made from discrete covalent molecules and have both intramolecular and intermolecular forces at work.

Review Questions page 282 - 283: 7, 9, 10, 14, 15, 16, 17, 18, 19, 20, 21, 22, 23 a) only, 24, 25.

Detailed Classification Scheme

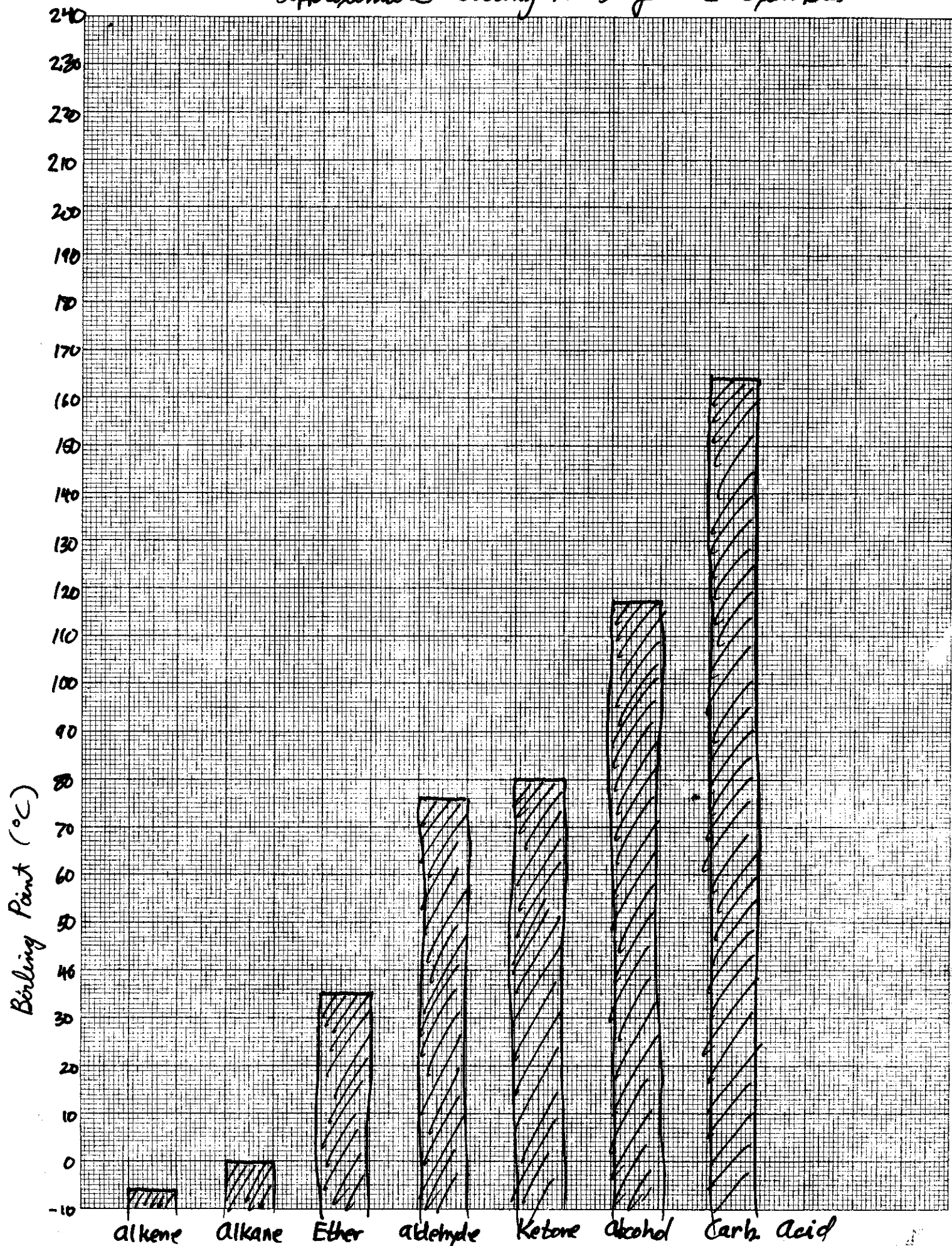
Molecule Size	Type of Compound	Primary Force Intramolecular Force	Secondary Force Intermolecular Force	Examples
<p>Macromolecules</p>	<p>Covalent Network Solid "Network Solid" (Three Dimensional)</p>	<p>Covalent (Intramolecular)</p>	<p>N.A.</p>	<p>diamond (C_n) quartz (SiO₂)</p>
	<p>Covalent Network Solid (Two Dimensional)</p>	<p>Covalent (Within Planes - Intraplantar)</p>	<p>van der Waals (Between Planes - Interplanar)</p>	<p>graphite (C_n)</p>
<p>Discrete Covalent Molecules</p>	<p>Ionic Solid (Pure Ionic)</p>	<p>Ionic (Intramolecular)</p>	<p>N.A.</p>	<p>NaCl</p>
	<p>Ionic Solid (Polyatomic Covalent Ions)</p>	<p>Covalent Within Ions (Intraionic)</p>	<p>Ionic (Between Ions - Interionic)</p>	<p>NH₄NO₃</p>
<p>Discrete Covalent Molecules</p>	<p>Metallic Solid</p>	<p>Metallic (Intramolecular)</p>	<p>N.A.</p>	<p>Fe</p>
	<p>Non-polar Molecules</p>	<p>Covalent (Intramolecular)</p>	<p>Hydrogen Bond</p>	<p>H₂O</p>
<p>Discrete Covalent Molecules</p>	<p>Polar Molecules</p>	<p>Covalent (Intramolecular)</p>	<p>Dipole</p>	<p>CH₃F</p>
<p>Discrete Covalent Molecules</p>	<p>Polar Molecules</p>	<p>Covalent (Intramolecular)</p>	<p>van der Waals</p>	<p>C₆H₁₂</p>

Types of Solids

COVALENT NETWORK	 <p style="text-align: center;">Diamond Graphite</p> <p style="text-align: center;">Silica</p>	<p>intramolecular* → covalent intermolecular* → N.A. *graphite covalent intraplanar v.d.W interplanar</p> <p>lattice points: occupied by neutral atoms (C or Si and O)</p>
		<p>conductivity*: N.A. (atoms neutral, electron locked in bonds) * graphite resonant pi electrons allow for conductivity through planes, no conductivity between planes (anisotropic)</p>
		<p>solubility: N.A. (replacement force for a covalent bond is next to impossible)</p>
IONIC	 <p style="text-align: center;">SODIUM CHLORIDE LATTICE STRUCTURE</p>	<p>intramolecular → ionic bond intermolecular → N.A. lattice points: occupied by ions (Na^+ Cl^-) {if NH_4NO_3 include an intraionic cov. bond}</p>
		<p>conductivity: (s) non-conductive (ions locked in lattice, electrons held tight in ions) (l) moderately conductive (Na^+ and Cl^- ions free to move) (aq) moderately conductive (hydrated Na^+ ions and Cl^- free to move)</p>
		<p>solubility in water: variable depending on hydration energy vs lattice energy (i.e. are there adequate replacement forces), units are individual hydrated ions (i.e. Na^+ & Cl^-) {if NH_4NO_3 units would be NH_4^+ and NO_3^-} solubility in non-polar: nil</p>
METALLIC	<p style="text-align: center;">Metallic Sea of Electrons</p>  <p style="text-align: center;">Electrons are not bonded to any particular atom and are free to move about in the solid.</p>	<p>intramolecular → metallic bond intermolecular → N.A. lattice points: occupied by individual metallic ions</p>
		<p>conductivity: excellent in (s) and (l) states due to highly mobile free moving electrons</p>
		<p>solubility: soluble in other metals only once warmed past melting point (an alloy is a metallic solution of more than one type of metal - also see mercury amalgam, units are individual metal ions)</p>
MOLECULAR		<p>intramolecular → covalent bond intermolecular → depends on molecules (ice has an H-bond, I_2 would be v.d.W only) lattice points: occupied by individual molecules</p>
		<p>conductivity: nil in solid state, if molecule is polar and capable of ion dissociation, slight conductivity is observed in the liquid state</p>
		<p>solubility: if replacement forces are adequate (similar polarity and similar new intermolecular forces) can be highly soluble, units will be individual molecules (molecules do not fall apart due to covalent bond framework)</p>

	ALKANES	ALKENES	ALCOHOLS	ALDEHYDES	KETONES	CARB. ACIDS	ETHERS
1	methane CH ₄ m.p. -183 d.p. -162		methyl alcohol -CH m.p. -97 d.p. 65	formaldehyde m.p. -92 d.p. -20		formic acid m.p. -8 d.p. 107	
2	ethane - m.p. -183 d.p. -89	ethene = m.p. -169 d.p. -104	ethyl alcohol m.p. -115 d.p. 78.5	acetaldehyde m.p. -121 d.p. 20		acetic acid m.p. -17 d.p. 118	diethyl ether m.p. -140 d.p. -24
3	propane m.p. -182 d.p. -42	propene m.p. -185 d.p. -47	n-propyl alcohol m.p. -127 d.p. 97	propanal m.p. -81 d.p. 49	dimethyl ketone (acetone) m.p. -94 d.p. 56	propanoic acid m.p. -2 d.p. 144	
4	butane m.p. -138 d.p. -0.5	1-butene m.p. -165 d.p. -6	n-butyl alcohol m.p. -90 d.p. 117	butanal m.p. -99 d.p. 76	ethyl methyl ketone m.p. -86 d.p. 80	butanoic acid m.p. -6 d.p. 164	diethyl ether m.p. -116 d.p. 35
5	pentane m.p. -129 d.p. 36	1-pentene m.p. -138 d.p. 30	n-pentyl alcohol m.p. -79 d.p. 138	pentanal m.p. -91 d.p. 133	diethyl ketone m.p. -48 d.p. 103	pentanoic acid m.p. -34 d.p. 186	
6	hexane m.p. -95 d.p. 69	1-hexene m.p. -140 d.p. 63	n-hexyl alcohol m.p. -47 d.p. 158	hexanal m.p. -56 d.p. 128		hexanoic acid m.p. -3 d.p. 205	diisopropyl ether m.p. -60 d.p. 69
7	heptane m.p. -91 d.p. 98	1-heptene m.p. -119 d.p. 94	n-heptyl alcohol m.p. -34 d.p. 176	heptanal m.p. -42 d.p. 154	diisopropyl ketone m.p. -85 d.p. 151	heptanoic acid m.p. -8 d.p. 223	
8	octane m.p. -57 d.p. 126	1-octene m.p. -102 d.p. 121	n-octyl alcohol m.p. -17 d.p. 195	octanal m.p. 171		octanoic acid m.p. 11 d.p. 239	di-t-butyl ether m.p. 70 d.p.

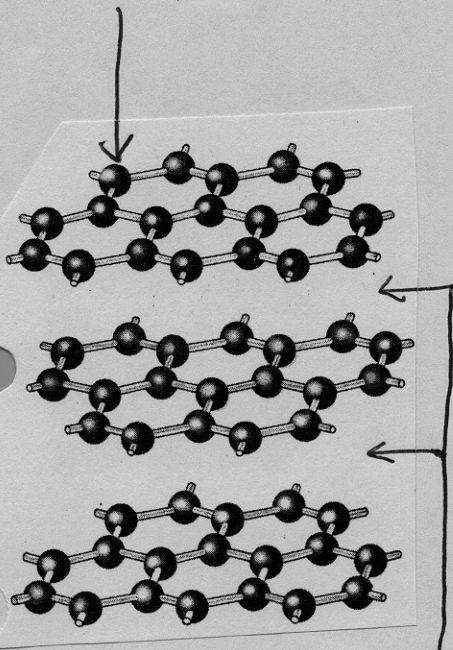
Approximate Boiling Points of 4C Examples



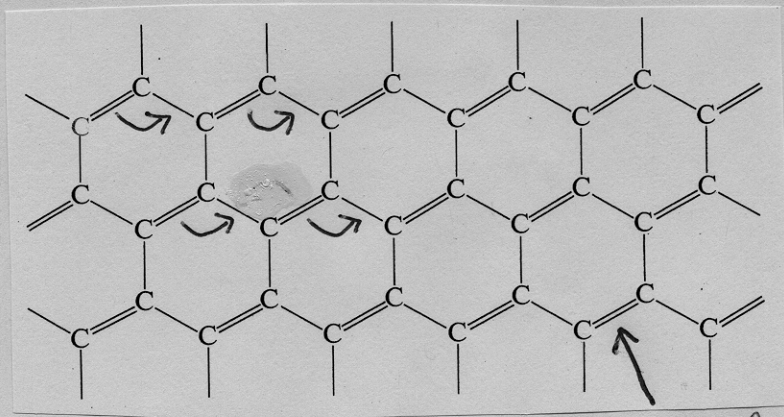
Graphite

Resonance Structures of Graphite

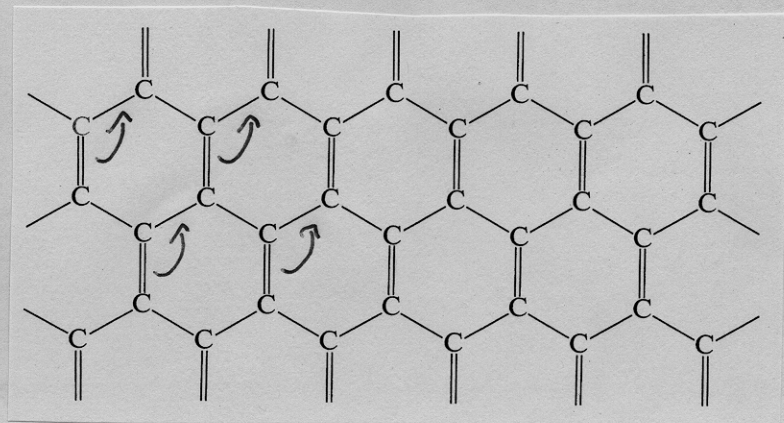
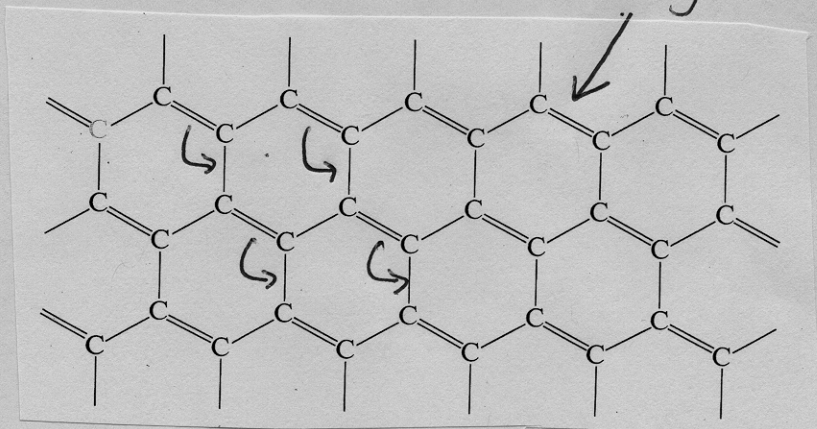
σ bonded planes
with delocalized
 e^- in π bonds
(not shown)



van der Waals forces
between planes



delocalized
 π bonding e^-



Which Type of Bond?

Elements that form discrete covalent molecules or monatomic "molecules" (M.P. and B.P. is governed by intermolecular van der Waal forces only)

H₂																	He		
																N₂	O₂	F₂	Ne
																P₄	S₈	Cl₂	Ar
																		Br₂	Kr
																		I₂	Xe
																		At₂	Rn

Elements that are capable of creating covalent network solids (add (SiO₂)_n to this list) (M.P. and B.P. is governed by intramolecular covalent bonds within a macromolecule and as such is very high)

H₂																	He				
																B	C	N₂	O₂	F₂	Ne
																		P_n	S₈	Cl₂	Ar
																Ge	As	Se	Br₂	Kr	
																Sn	Sb	Te	I₂	Xe	
																		Bi	Po	At₂	Rn

The elements with the lightest shading are molecular liquids and solids bound by van der Waals forces. The elements with medium shading are covalently bonded network solids. The metallic elements are indicated by the darkest shading.

H₂																	He
Li	Be											B	C	N₂	O₂	F₂	Ne
Na	Mg											Al	Si	P_n*	S₈	Cl₂	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br₂	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I₂	Xe
Cs	Ba	La-Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At₂	Rn
Fr	Ra	Ac-Lr															
		La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	
		Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr	

* Phosphorus has several different molecular forms. Volatile white phosphorus is P₄. Non-volatile, insoluble red phosphorus is polymeric P_n.

TYPES OF COMPOUNDS

	IONIC (macromolecules)	COVALENT			METALLIC (macromolecules)
		NETWORK SOLID (macromolecules)	DISCRETE COVALENT MOLECULES		
			POLAR	NON-POLAR	
STRUCTURE	a very large number of positive metallic ions and negative non-metallic ions formed by a transfer of electrons, organized into a crystal lattice structure, held together by strong electrostatic forces of attraction between oppositely charged ions, ions obey the octet rule, entire crystal is electrically neutral, macromolecule	many neutral non-metallic atoms, organized into a crystal lattice structure, held together by a network of strong covalent bonds, bonding obeys the octet rule, macromolecule	a specific number of non-metallic atoms bonded together by covalent bonds, any bond polarizations will add to give a net polarization with regions of partial negative and partial positive charge, discrete molecule	a specific number of non-metallic atoms bonded together by covalent bonds, any bond polarizations will cancel out giving no net polarization, discrete molecule	many positively charged metallic ions surrounded by a sea of free moving electrons, metal ions are packed together as close as possible, held together by strong electrostatic forces of attraction between positive ions and free moving electrons, macromolecule
EXAMPLES	KCl, CaF ₂ , Cs ₂ S, MgO, LiF, NaCl (empirical formula)	C _(n) - diamond, SiO ₂ - quartz SiC, Si (empirical formula)	H ₂ O, ICl, NH ₃ , CH ₃ COOH (molecular formula)	CH ₄ , C ₄ H ₁₀ , wax, O ₂ , N ₂ , Cl ₂ (molecular formula)	Cu, Au, Ag, Na, U, and many alloys
MELTING AND BOILING POINTS	high, strong intramolecular forces must be overcome for melting/boiling to occur**	high, strong intramolecular forces must be overcome for melting/boiling to occur**	low to moderate, due to weaker intermolecular electrostatic interactions*	low due to very weak intermolecular interactions*	usually high, strong intramolecular forces must be overcome for melting/boiling to occur**
PHYSICAL APPEARANCE AND RESPONSE TO PHYSICAL STRESS	well defined solid crystals, will shatter or cleave*** along a plane surface within the crystal lattice	well defined solid crystals, will shatter or cleave*** along a plane surface within the crystal lattice	often liquids or gases at room temperature, solid may be crystalline, often crumbly when stressed	often liquids or gases at room temperature, solids will be soft and mushy like wax or bacon grease	high luster solids, malleable and ductile when stressed, can be cast into forms, positive metallic ions can carefully slip past each other when stressed
SOLUBILITY	insoluble in non-polar solvents, variable solubility in polar solvents such as water (lattice**** energy versus hydration**** energy)	insoluble in all solvents, neutral atoms do not have a solvent/solute interaction capability	insoluble in non-polar solvents, soluble in polar solvents (like dissolve like), good solvent/solute interactions between δ ⁺ and δ ⁻	soluble in non-polar solvents (like dissolves like), insoluble in polar solvents	insoluble, electrons cannot dissolve because they move to fast (soluble in liquid mercury - like dissolves like)
CONDUCTIVITY	solid - non conductive, liquid (molten) state - conductive (free moving ions), aqueous solution - conductive (free moving ions)	non conductive in any state (no charged particles)	non conductive in any state (no particles that carry a net charge)	non conductive in any state (no charged particles)	conductive in solid and liquid state due to free moving electrons

* when melting/boiling a discrete covalent molecules only intermolecular attractions need be overcome to create a particle that is small enough to behave as a liquid or a gas

** when melting/boiling a macromolecule the intramolecular forces between atoms/ions must be overcome to create particles small enough to form a gas or a liquid, since intramolecular forces are 10 to 100 times stronger than intermolecular forces, macromolecules have a much higher melting/boiling point than discrete covalent molecules

*** cleavage is the property exhibited by a substance when it can fracture along a plane of particles in the crystal lattice resulting in smooth plane surfaces

**** lattice energy is the energy that corresponds to bonding within a lattice, hydration energy is the energy that corresponds to attraction between ions and water molecules