Unit 2 – Chemistry 2202
CHEMICAL BONDING IN MATTER (MHR Text p. 162 - 164)
Chemical Bonds: are attractive electrostatic forces that hold atoms or ions together in a substance.

Bohr's Model of the Atom
- protons and neutrons are located in the nucleus
- electrons are located in orbits or specific energy levels outside the nucleus.

Bohr's Energy Level Diagrams
- # of e^- in an atom is equal to its atomic number
- # of energy levels is equal to the period number
- Maximum number of electrons in the first three energy levels is 2, 8, 8
- # of electrons in the outermost energy level is equal to the group number of the element.
- energy level diagrams for ions resemble the diagrams for the nearest noble gas (full outer energy level; therefore stable and unreactive)

Examples: Draw Bohr diagrams and energy level diagrams for the following atoms: Sodium ion, magnesium atom, nitride ion and chlorine.

Valence Energy Level: the outermost energy level of an atom, which contains electrons.

Valence Electrons: the electrons in the highest occupied energy level of an atom.

Note: The number of valence electrons of an element (in Group A) is equal to the group number in which it is found.

Group #	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIA
# of Valence Electrons	1	2	3	4	5	6	7	8(except He)

Lewis Structures or Electron Dot Diagrams
When two atoms approach each other, it is their outermost energy levels (valence shells) that come in contact first.

Gilbert N. Lewis: To show formation of chemical bonds we need only show valence electrons rather than all electrons as in Bohr Energy Level Diagrams.

Rule: For each element,
a) write down the symbol of the element. This represents the nucleus and inner energy level electrons.
b) use the periodic table to determine the number of valence electrons for the atom.
c) Place the dots (representing electrons) in four areas (orbitals) around the atom. **DO NOT** form pairs, until each of the four areas has at least one dot.

Examples:

Symbol	Dot Diagram	# lone pairs	# bonding pairs	bonding capacity
Li				
Be				
B				
C				
N				
O				
F				
Ne				

Note: In Lewis Structures electrons exist as single electrons or as paired electrons.

Bonding Electron (or 'unpaired electron') is a single electron in a valence orbital of an atom. Bonding electrons are always involved in bonding.

Lone Pairs (non-bonding electrons) are electrons found as pairs in filled valence orbitals of an atom. They are rarely involved in bonding.

Bonding Capacity is the maximum number of bonds an atom can form. It equals the number of bonding electrons in the atoms.

Bonding in a Molecular Compound (MHR Text p. 168 - 171)

Covalent Bond: the force of attraction that exists between two nonmetallic atoms whose nuclei **share** an attraction for the same pair of electrons.

Lewis Dot Diagrams for Molecular Compounds

1) Draw a dot diagram for each element.
2) Place the atom with the most bonding electrons is placed in the centre with the other atoms bonded to it.
3) Count the number of bonding electrons for each atom and pair bonding electrons between atoms to form single, double or triple bonds. **NO electrons are to be left unpaired.**
4) Count to ensure each atom has an octet (except hydrogen & boron)

Molecular Elements two or more covalently bonded atoms of one type of element.

Ex: Draw Lewis Dot Diagrams for H_2, Cl_2 , O_2, N_2

A Single bond is when there is **ONE** electron pair shared between two atoms.
A Double bond is when there are **TWO** electron pairs shared between two atoms.

A Triple bond is when there are **THREE** electron pairs shared between two atoms.

Molecular Compounds contain two or more different kinds of non-metallic atoms covalently bonded together.

Structural Formulas are a simplified way to show bonding present in a molecule.

1. a dash (-) is used to represent a bonding pair of electrons.
2. lone pairs are omitted.

Ex: Draw Lewis Dot Diagrams and Structural Diagrams for H_2O, CO_2, CH_4 , NH_3 , CH_2O , C_3H_8 , C_2H_4 , C_2H_2 , HCN, and CH_3OH. **Place in your Notebook**

Molecular Shapes *(MHR Text p. 189 - 193)*
All molecules have a definite 3-D shape.

Stereochemistry is a study of the shape of chemical species.

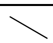
VSEPR Theory (Valence Shell Electron Pair Repulsion)

- Used to predict molecular shape (geometry).
- Developed by Canadian chemist Ronald Gillespie.

Molecular geometry can be determined from the Lewis Dot Diagram.

Rules for VSEPR Theory

1. Valence electron pairs, both bonding (shared) and lone pairs, arrange themselves around the central atom in the molecule so as to minimize repulsion. Therefore, the electrons are situated as far away from one another as possible.
2. Multiple bonds are treated as single bonds.
3. Actual shape will depend on how many pairs are bonding and how many are lone pairs.
4.

	atom is in same plane as central atom
	atom goes behind the plane (away from you)
	atom goes in front of plane (towards you)

Five Main Shapes
(There are others not dealt with in this course)

1. **Tetrahedral**

Molecule	Lewis Dot	Around Central Atom	Shape & Bond Angles
CH_4			

2. **Pyramidal**

NBr_3			

3. **V-shaped or Bent**

Molecule	Lewis Dot	Around Central Atom	Shape & Bond Angles
H_2O			

4. **Trigonal Planar**

BF_3			
CH_2O			
C_2H_4			

5. **Linear**

C_2H_2			

Practice: For each of the following draw the Lewis Dot Diagram, draw the molecular shape; indicate the name of the shape and bond angles.

SiH_2Cl_2	PH_2Br	SI_2	$C_2H_2Cl_2$
C_2HBr	HCl	HCN	CH_3OH

<u>Ionic compounds</u>

-Are formed when metals lose electrons (transferring electrons) to the nonmetals, which combine to form ionic bonds

Cations: - are formed when metals *lose* one or more valence electrons.

Anions: - are formed when nonmetals *gain* one or more valence electrons.

Draw Lewis dot diagrams for ions and Ionic compounds

A) Sodium chloride B) calcium chloride C) magnesium oxide

Electronegativity (MHR Text p. 174 -175)

Draw an electron dot formula for HCl . ________

The formula suggests that the pair of electrons, which constitutes a covalent bond, is shared equally between the hydrogen and chlorine.

This is not the case. In hydrogen chloride, the chlorine atom exercises a stronger attractive force on the bonding pair than the hydrogen atom does. We say chlorine is more <u>electronegative</u> than hydrogen.

<u>Electronegativity</u> (EN) is the tendency of an atom to attract electrons to itself when it is chemically combined with another atom.

Linus Pauling proposed numerical values for the electronegativities of the elements which are numerical value locate on your periodic table.

Trends in Electronegativity

1. Increase in EN from left to right within a period.

2. Decrease in EN from top to bottom within a group.

*The two trends combine to give fluorine the highest EN at 4.0 and cesium the lowest EN at 0.7.

Notes:
1) **EN of metals is low.**
Metals have very little attraction for their own electrons or those of other atoms. They hold their valence electrons very loosely and will lose them easily to other atoms.

2) **EN of nonmetals is high**.
Nonmetals have a strong attraction for their own electrons and those of other atoms. They hold their valence electrons very strongly and will even gain more from other atoms.

Polarity (MHR Text p.176-178)

Polarity: The differences in the EN of atoms given two different types of covalent bonds.

Non-polar Covalent Bond:occurs when both atoms involved in bonding have the same EN. The bonding electron pair is shared equally and is uniformly found between the nuclei of two atoms.

EX: 1) H - H 2) Cl - Cl 3) -S - I
2.1 2.1 3.0 3.0 2.5 2.5

Polar Covalent Bond:occurs when the two atoms involved in bonding have different EN. The bonding electron pair is unequally shared and therefore the electrons are not symmetrically distributed (equally spread) between the two atoms. There is an "electron shift".
EX: 1) $\rightarrow$
C - O
2.5 3.0
δ^+ δ^-

Notes:
1) The more electronegative atom will become partially negative (δ^-). Since the shared electron is pulled closer, it appears to have more electrons surrounding it than it has protons in its nucleus.

2) The least electronegative atom will become partially positive (δ^+). Since its shared electron is pulled slightly away, it appears to have less electrons surrounding it than it has protons in its nucleus.

3) These electrons are stilled shared, they are NOT exchanged - only shifted.

4) Polar Covalent Bonds have bond dipoles (a δ^+ end and a δ^+ end) but are still electrically neutral.

5) Bond dipoles are represented by an arrow with the arrowhead pointing towards the partially negative site.

Molecular Polarity (MHR Text p. 194 - 196)

Note:
1) A molecule is non-polar, if all bond dipoles cancel.
(central atom does not move, if you look at the bond dipoles as pushing or pulling on the central atom).

2) A molecule is polar if bond dipoles do not cancel. (central atom does move)

Examples: What is the bond polarity and the molecular polarity of the compounds? **Place in your Notebook**

CH_4, CF_4 , CH_3Br , CH_3I, CI_4

NH_3 , NF_3, NCl_3

H_2O, OF_2, SI_2

BF_3, BH_2Cl, CH_2O, CI_2S, C_2H_3Cl, $C_2H_2F_2$ (exceptions)

C_2H_2 , C_2F_2, C_2HF, HCN, HCl, CO_2

General Rules:

1) All molecules will be **<u>nonpolar</u>** if, all atoms including the central atom have the same electronegativity (no bond dipoles set up). Ex: CI_4, NCl_3, SI_2, CI_2S, CS_2

2) **Tetrahedral, trigonal planar , and linear molecules** will be:
<u>non-polar</u>: if all atoms surrounding the central atom are the same. (bond dipoles cancel)
<u>polar</u>: if one (or more) atom(s) surrounding the central atom are different. (bond dipoles do not cancel)

3) **Pyramidal and v-shaped molecules** will be **<u>polar</u>**. (bond dipoles do not cancel)

4a) **Hydrocarbons** (compounds made up of hydrogen and carbon atoms only, C_xH_y) are **<u>non-polar</u>.**
Ex: CH_4, C_2H_4, C_2H_2
b) Hydrocarbons become **<u>polar</u>** if one (or more) atom(s) is different. ie: $C_xH_yA_z$
Ex: CH_3Br, C_2H_3Cl, CH_3OH, C_2HCl

Question. Which of the following molecules are polar?
Answer. Λ if you know the shape, you can use the general rules.
Λ to determine the shape look at the first atom in a formula; it is ***<u>usually</u>*** the central atom. Look at the group it is in to determine # of bonding and lone pairs. (For C and Si in particular, pay attention to how many things are bonded to it)
EX: CH_4 (group IV - 0 LP, 4 BP tetrahedral)
CH_2O (group IV - 0 LP, 3 BP tri. planar)

Molecule	Shape & Polarity	Molecule	Shape & Polarity
CH_4		SiH_3Cl	
NCl_3		PBr_3	
OF_2		SI_2	
BH_2Cl		C_2H_4	
Si_2Cl_2		HCN	

Solubility (MHR Text p. 247)

Solubility: refers to the ability of a substance to dissolve in a given solvent.
General Rule: like dissolves like
thus: 1) polar dissolves polar
2) non-polar dissolves non-polar
3) polar and non-polar do not dissolve.

Question. From the above question, which of the compounds will dissolve in water?

Answer. - When looking at a solubility question, you need to determine **polarity** of the compounds. Water is polar. (It is a v-shaped molecule). Since like dissolves like, it will only dissolve other polar molecules.

– From the above question, water will dissolve SiH_3Cl, PBr_3, OF_2, BH_2Cl, HCN.

Question. Why do water and oil (a hydrocarbon) not mix?

Answer. The general rule is like dissolves like. Water is a polar molecule whereas oil is a nonpolar one since it is a hydrocarbon. Therefore, the two will not mix.

Intermolecular Forces (MHR Text p. 202 - 206)

Molecular compounds are:

1) made up of molecules
2) are made up of two or more nonmetallic atoms
3) held together by covalent bonds (sharing e^-)

Q. But what holds the individual molecules together in a compound?

A. **Intermolecular Forces**:are the weak forces of attraction which hold molecules together in a compound.

NOTE:

The presence of intermolecular forces can be used to explain observations, such as surface tension, change in state, and boiling point.

Q. What happens when you boil (melt) a compound?

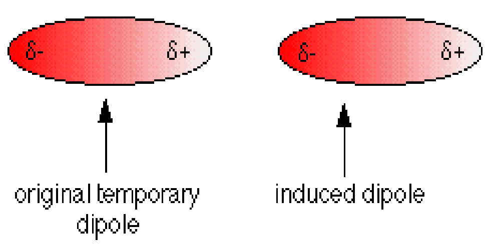
A. In going from a solid to a liquid to a gas, the molecules are moved further apart. To do this the forces which hold the molecules together must be broken. The stronger the intermolecular forces, the higher the boiling (melting) point will be.

Two Types of Intermolecular Forces

1. van der Waals Forces
   a) London Dispersion Forces (LDF)
   b) dipole-dipole (D-D)
2. Hydrogen Bonding (H- Bond)

London Dispersion Forces

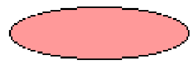
- force of attraction between two or more molecules as a result of electron movement setting up temporary fluctuating dipoles; these dipoles ***induce*** opposite charge dipole in a neighbouring molecule; thus attract



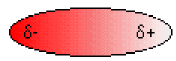
- are the main force of attraction between non-polar molecules
- are intermolecular forces that exists between all molecules.

Explaining London Dispersion Forces
Temporary fluctuating dipoles

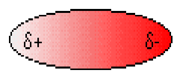
The diagram represents a small symmetrical molecule - Br_2. The even shading shows that on average there is no electrical distortion



But the electrons are mobile, and at any one instant they might find themselves towards one end of the molecule, making that end δ^-. The other end will be temporarily short of electrons and so becomes δ^+.



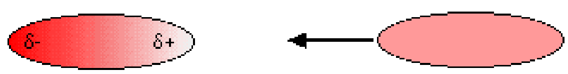
An instant later the electrons may well have moved up to the other end, reversing the polarity of the molecule.



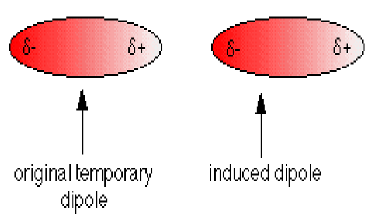
This constant "sloshing around" of the electrons in the molecule causes rapidly fluctuating dipoles even in the most symmetrical molecule. It even happens in monatomic molecules - molecules of noble gases.

Temporary dipoles give rise to intermolecular attractions

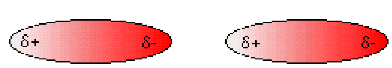
Imagine a molecule, which has a temporary polarity being approached by one which happens to be entirely non-polar just at that moment.



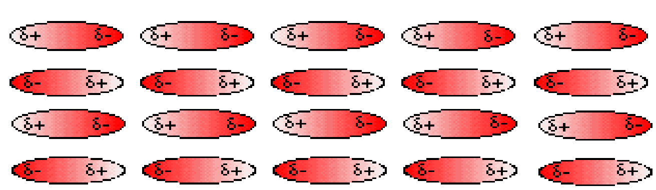
As the right hand molecule approaches, its electrons will tend to be attracted by the slightly positive end of the left hand one. This sets up an ***induced dipole*** in the approaching molecule, which is orientated in such a way that the δ^+ end of one is attracted to the δ^- end of the other.



An instant later the electrons in the left hand molecule may well have moved up the other end. In doing so, they will repel the electrons in the right hand. The polarity of both molecules reverses, but you still have δ^+ attracting δ^-. As long as the molecules stay close to each other the polarities will continue to fluctuate in synchronisation so that the attraction is always maintained.



There is no reason why this has to be restricted to two molecules. As long as the molecules are close together this synchronised movement of the electrons can occur over huge numbers of molecules.



This diagram shows how a whole lattice of molecules could be held together in a solid using van der Waals dispersion forces. An instant later, of course, you would have to draw a quite different.

Factors Affecting the Strength of LDF

1) # of electrons in a molecule - the more electrons a substance has, the more easily an uneven distribution of charge can occur, the stronger the LDF.

2) shape of the molecule - the more spherical the shape, the less surface area there is , thus less opportunity to induce a charge on a nearby molecule; the weaker the LDF

Q. Compare the boiling points of the nonpolar halogens with the number of electrons each has.
Answer.

Halogen	Boiling Point (°C)	# of Electrons
fluorine, F_2	-188	
chlorine, Cl_2	-34.6	
bromine, Br_2	58.8	
iodine, I_2	184	

The trend is confirmed. We would predict that as the number of electrons increases, the London dispersion forces would become stronger, thus require more energy to break these forces - ie: have a higher boiling point.

Q. Butane, C_4H_{10}, has a much higher boiling point than Cl_2, thus must have stronger intermolecular forces. Give a reason for the difference in boiling points.

Note: Molecules that have the same number of electrons are said to be **<u>isoelectronic</u>**.

Problem: CO_2 and Cl_2 contain only London dispersion forces. Explain which of these compounds will have the higher boiling point.

Dipole-Dipole Forces: is the simultaneous attraction of a molecular dipole by the surrounding molecular dipoles of opposite charge. (therefore occurs in polar molecules only)

```
   --->        --->
 H -- Cl     H -- Cl
 +     -     +     -
   <---        <---
 Cl -- H     Cl-- H
 -     +     -    +
   --->        --->
 H -- Cl     H -- Cl
 +     -     +     -
```

- the partially positive end of one molecule attracts the partially negative end of a neighbouring molecule.
- the strength of these forces will depend on the difference in the electronegativity of the atoms bonded together.

Question 1: List the hydrogen halides in increasing strength of dipole-dipole forces.

H-F, H-Cl, H-Br, HI

Question 2: Explain which of the following compounds, F_2 or CH_3F, has the higher boiling point.

Question 3: Which of the following substances, ICl or CCl_4, has the higher boiling point?

Question 4: Which of the following substances, I_2 or IBr, has the higher boiling point?

Note: LDF are a stronger force than dipole-dipole.

Question 5: Predict the boiling point of HF given the following data.

Hydrogen Halide	Number of Electrons	Boiling Point (EC)
HF	10	?
HCl	18	-83.7
HBr	36	-67.0
HI	54	-35.4

Answer: Based on the number of electrons, strength of LDF, one would predict that HF has a lower boiling point than -83.7EC. It in fact has the highest boiling point of all at 19.4EC.

Note:

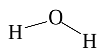
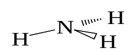
1) Increasing strength of LDF with increase in number of electrons holds true when comparing HCl, HBr, HI.

(Question: WHY NOT WITH HF?)

2) While HF is the most polar, molecular polarity alone could not account for the magnitude of reversal in trend.

3) Suggests the existence of an additional intermolecular force greater than van der Waals forces. --- Hydrogen Bonding.

Question. So, what do,

H-F  

have in common?

Answer: They each contain a highly polar bond O - H, N - H, H - F, with the positive end being hydrogen.

Hydrogen Bonding: simultaneous attraction of a proton (very partially positive hydrogen) by the electron pairs of adjacent N, O, or F atoms

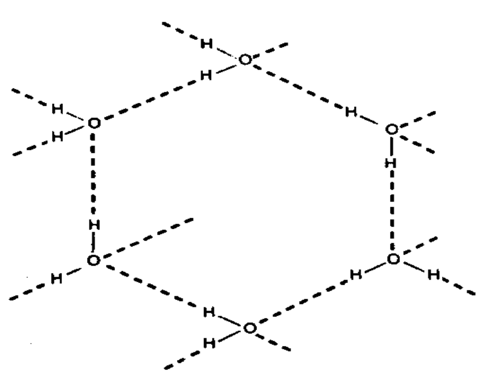
This type of bonding arises when a molecule contains a highly polar bond; O-H, N-H, or H-F, (less commonly H-Cl); the positive end being hydrogen.

The charge separation (shared electrons are closer to the more electronegative atom) is great enough that the H atom is attracted to a highly electronegative atom in another molecule having unshared electron pairs.

Example: Hydrogen fluoride.

```
  ..        ..        ..
H-  :F: --- H-  :F: --- H-  :F:
```

Hydrogen bonding is represented using dotted lines (represents a weaker bond type than a covalent bond).
Example: Water (ICE)
Due to H-bonding & v-shape
1) snowflakes have 6 sides
2) holes with air that why ice floats



Uniqueness of Hydrogen
Hydrogen has no other electrons other than the one involve in the covalent bond linking it in the molecule. Thus, the hydrogen atom's positively charged nucleus can interact easily with unshared electrons on a neighbouring molecule.

Question. Which of the following compounds exhibit H-bonding: H_2O, CH_4, CHF_3, NF_3, CH_3OH?
Answer. Draw structural diagrams for each.

H_2O	CH_4	CHF_3	NF_3	CH_3OH

Properties of Hydrogen Bonded Substances

Hydrogen bonding affects both physical and chemical properties of substances.

1. increased melting and boiling points

Pair	Substance	Boiling Point (℃)	# of electrons	hydrogen bonding?
1	HF	19.4	10	yes
	HCl	-83.7	18	no
2	H_2O	100	10	yes
	H_2S	-61	18	no
3	NH_3	-40	10	yes
	PH_3	-90	18	no

2. increased solubility of substances mutually involving hydrogen bonding.

EX: water, H_2O; methanol, CH_3OH; and ethanol, C_2H_5OH are soluble in each other in all proportions. They all exhibit H-bonding.

3. shape and stability of certain structures.

1) when water freezes hydrogen bonding directed by the v-shape of water leave hexagonal holes. This structure is less dense than water -- it would not be possible without hydrogen bonding.

2) the behaviour and stability of protein molecules within the body, depends on their shape which is affected by the hydrogen bonding present.

Question 1 : List the following substances, C_3H_8, CH_3Cl, C_2H_5OH in order of increasing boiling point? Give reasons.

Question 2 : Which substance, C_4H_{10} or C_2H_5OH has the highest boiling? Give reasons for your answer.

Question 3: Why is it difficult to determine whether C_5H_{12} or CH_3OH has the higher boiling point?

Aside: One would always ***predict*** the H-bonded substance as having the higher boiling point.

Intramolecular Forces
Ionic compounds, molecular compounds and metals are distinguishable based on some very different properties. These differences exist because of the type of bonding that holds the different types of compounds together.

Intramolecular Forces: are the forces of attraction or bonds that exist between the atoms or ions of a compound (now we are talking about inside the molecule, whereas intermolecular was outside).
Ex: 1) covalent bonds
2) ionic bonds
3) metallic bonds

Type of Compound and Type of Bond (MHR Text P. 176 -178)
1. **Ionic compounds** are held together by **ionic bonds** between metals and nonmetals.
– Metals (low EN elements) **lose (transfer)** their valence electrons to nonmetals (high EN elements).
2. **Molecular compounds** are held together by **covalent bonds**.
– Covalent bonds involve a **sharing** of valence electrons between nonmetallic elements (high EN).
3. **Metals** are held together by **metallic bonding**.
– Metallic bonds form between atoms which have low EN.

Ex: X (3.2), A (0.9), Q (1.2), M(2.8) are fictitious elements with the given EN. Predict the type of bonds found between the pairs of atoms listed.
a) X & M b) A & Q c) A & X d) X & X

Note: EN differences result in bonds having a percent of ionic and covalent character (Table 5.1 p. 178).
ex: Since polar covalent bonds (ex: X & M) have some similarity with an ionic bond we say it has percent ionic character.

To determine the type of bond, look at difference in EN

- 0 is non-polar
- 0 -0.5 is slightly polar
- 0.5 - 1.7 is mostly polar covalent
- 1.7 above is mostly ionic bonds

Ionic Bonding (MHR Text p. 165-168)

Simple ions form because atoms lose or gain electrons to achieve an e^- configuration like the nearest noble gas.

1) metals lose electrons - form +ve ions
2) nonmetals gain electrons - form -ve ions

Example:

transfer

Na	F
$1e^-$	
$8e^-$	$7e^-$
$2e^-$	$2e^-$

Na • :F: (Lewis)

Bohr becomes Lewis

$11p^+$	$9p^+$
$8e^-$	$8e^-$
$2e^-$	$2e^-$

$[Na]^+ [:\ddot{F}:]^-$

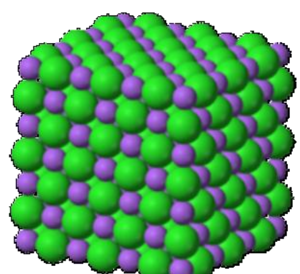
Na^+<-ionic bond-> F^- (see p. 166)

Ionic compound is a compound formed when positive ions (metallic) combine with negative ions (nonmetallic). They are held together by ionic bonds.

Ionic Bond is the simultaneous attraction of an ion by its surrounding ions of opposite charge within an ionic crystal.

Crystal is a 3-D, continuous, repeating pattern of positive and negative ions in an ionic solid.

Ex: NaCl (cubic structure)



each Na^+ is surrounded by 6 Cl^-

each Cl^- is surrounded by 6 Na^+

Formula: Na_6Cl_6 Smallest Unit NaCl each Na+ is surrounded by 6 Cl-

Formula Unit is the smallest possible unit in an ionic crystal.

Properties of Ionic Compounds

1. solids at room temperature.
2. hard and brittle.
2. high melting and boiling points.
3. usually high solubility in H_2O
4. good conductors in liquid state and aqueous sol^n; poor conductor in solid state.

Aqueous - means dissolved in water.

Explaining Ionic Properties

1. **State, Hardness, Melting Point**

To break apart an ionic compound, you have to separate the individual positive and negative ions. Each ion has several strong ionic bonds to its neighbours - therefore it will require a large amount of energy to break all these bonds.

1) physical energy is not enough to break the bonds which accounts for its hardness and state.
2) large amount of heat energy needed to get ions moving fast enough to break the ionic bonds which accounts for the high melting point.

2. **High Solubility in Water**

Water is polar. The partially positive end will attract the negative ions in the ionic crystal, the partially negative end will attract the positive ions in the ionic crystal. The ionic crystal separates into ions in solution.

3. **Brittle (break easily when bent).**

If crystal is bent or hit, the positive and negative ions may be forced away from ions of opposite charge and closer to ions with the same charge. This shift results in repulsion and the crystal shatters.

push	strong repulsion
----->+ - + - + -	+ - + - + -
+ - + - + -	+ - + - + -
+ - + - + -	+ - + - + -
alternating pos. & neg. ions	crystal breaks

4. **Conductivity**

In order for a substance to conduct, there must be movement of charged particles (electrons, ions).

1) In solids, ions are not free to move because they are held in a fixed position by the ionic bonds.
2) In liquid state and aqueous solution, the ionic bonds have been broken and ions are now free to move towards an electrode of opposite charge.

Properties of Molecular Compounds

1. State:solids, liquids, or gases
2. Hardness: soft
3. Melting Point: low melting and boiling points.
4. Solubility in H_2O: varies from high to low
5. Conductivity: poor to none in all states.

Forces of Molecular Compounds

Intramolecular Force (inside of a molecule): covalent bonding

Intermolecular Force (between molecules): L.D.F., D-D, H-bonding

Explaining properties of molecular compounds

Q: Explain in terms of bonding present why molecular compounds are soft and have low melting points.

Answer: The individual molecules of a molecular compound are held together by weak intermolecular forces. Large amounts of energy are not needed to break these forces making them soft and causing them to have low melting points.

Q: Explain in terms of bonding present why molecular compounds are soluble in water and others are not.

Answer: Solubility in polar water will depend on the polarity of the molecular compound. If the intermolecular forces, dipole-dipole is present (molecule is polar), then the molecular compound will be soluble in water. If the intermolecular forces, dipole-dipole is not present (molecule is nonpolar), then it will not be soluble in water. Recall, polar dissolves in polar.

Q: Explain in terms of bonding present why molecular compounds do not conduct electricity in either state.

Answer: Molecular compounds do not conduct in either state because there are no charged particles available to move about. There are no ions present, and all electrons are held tightly by the very electronegative nonmetallic atoms or are involved in sharing for stability of the atoms.

Network covalent bonding

-Ionic compounds form crystals. These crystals are regular 3-D array of ions.

- So far we've looked at covalently bonded atoms, as they exist in individual molecules. However, it is also possible for covalently bonded compounds to exist in giant arrays.

Network Solid is a giant rigid structure in which nonmetallic atoms are covalently bonded together in a continuous 3-D array. The bonds are very strong and very directional in a 3-D network forming a single, very large molecule called a **macromolecule.** Ex 1) diamonds (C_n) 2) Silicon carbide(SiC) 3) Silicon dioxide (SiO_2)

Network covalent bonding is the simultaneous attraction of every atom to adjacent atoms by covalent bonds, within a 3-D lattice of atoms.

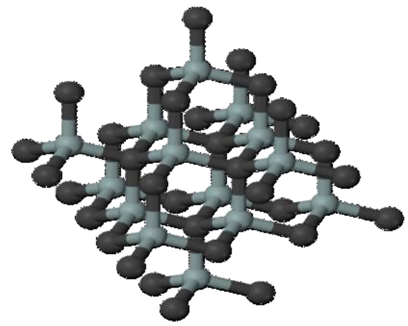
Properties of network solids

1) Strong and hard 2) high melting points 3) Not soluble on water 4) Do not conduct

Explaining properties of network covalent compounds

Q: Explain in terms of bonding present why network covalent solids are hard and have very high melting points.

Ans: Network covalent solids are hard and have high melting points due to each atom, within the network solid, being bound in 3 dimensions to four other atoms, and so on. This gives a tight 3-D network where each atom has a tetrahedral arrangement about it. All atoms are held in rigid positions, thus making it very difficult to break apart



Q: Explain, in terms of the bonding present, why network covalent solids do not conduct electricity.

Answer: Network covalent solids do not conduct electricity; due to the fact electrons are not free to move. All valence electrons within the network solid are involved in covalent bonds therefore cannot move about.

Metallic Bonding: found in metals - elements to the left of the staircase

$1e^-$

$2e^-$ $2e^-$ $2e^-$ $2e^-$

$2e^-$

$11p^+$

Na

Three things true for all metals

1) They have low electronegativity, so they do not hold on to their valence electrons strongly.
2) They have empty valence orbitals.
3) The loosely held valence electrons are free to move from one valence orbital to the next.

The atoms become positive ions because the valence electrons are dispersed. The positive ions remain fixed, but the electrons move and are simultaneously attracted to more than one nuclei.

A sample of a metal can be thought of as a group of cations in a "sea" of free moving valence electrons.

e^-
Na^+ Na^+ Na^+
e^- e^- e^-
Na^+ Na^+ Na^+ e^-
e^- e^- e^-
Na^+ Na^+ Na^+
e^-

Metallic Bond: is the simultaneous attraction of metal cations for the free moving valence electrons.

<u>Properties of Metals</u>

1)shiny, malleable, ductile 2) conducts in both solid and liquid 3) range of melting points

Explaining Properties of Metals

1. **Malleable and Ductile**.

The shifting of positive ions does not result in repulsion. The metallic bond is not fixed. As the metal ions move, the free moving valence electrons move too. The cations slide pass each other with the 'sea' of electrons acting as a lubricant.

```
----->+  e-  +  e-  +  e-                    +  e-   +  e-    +  e-
       e-       e-     e-                           e-        e-       e-
        +  e-   +  e-   +  e-          ----->       +  e-     +  e-     +  e-
          e-    e-     e-                          e-    e-     e-
           +     +      +                   +  e-     +  e-    +  e-
```

metal shape changes

2. **Good Conductors.**

   Electricity is the flow of electrons. When electrons are added from an outside source, they enter a substance full of free moving valence electrons. The valence electrons on the other end get pushed out to make room for the new electrons. This flow of electrons becomes an electric current.

3. **Metals have a Range of Melting Points.**

   The more valence electrons freely moving about, the larger the number of electrostatic attractions between stationary positive ions and valence electrons, the higher the melting point.

   - Sodium only has one valence electron/atom to contribute to the sea of electrons; melts at $97.8^{o}C$.

   -Magnesium has two valence electrons/atom to contribute to the sea of electrons; more attractions to overcome; melts at $649\ ^{o}C$.

4. **Metals form Alloys.**

   Alloys are solid solutions of two or more metals. The valence electrons of one metal are simultaneously attracted by another metal.

Order for Range of Strength of
Intra and Inter Molecular Forces

Generally:

Covalent (network)>Ionic (crystal lattice) > Metallic > H-bond > D-D > LDF

Exceptions LDF > D-D when there is more than $10e^-$

LDF > H-Bond when there is very complex molecule

ie. C_5H_{12} has 17 atoms compared to CH_3OH that has 6 atoms.

Q. List the following in order of decreasing boiling points.

Na, SiO_2 , C_3H_8 , $NaCl$, C_2H_5OH , OF_2

Answer Hint: What forces have to broken to boil them?

Na	metallic
SiO_2	network covalent
C_3H_8	LDF (26 e^-)
NaCl	ionic
C_2H_5OH	LDF (26 e^-), D-D, H-bond
OF_2	LDF (26 e^-), D-D

Decreasing order: $SiO_2 > NaCl > Na > C_2H_5OH > OF_2 > C_3H_8$