

SCH 4UI UNIT 1 Chapters 3 and 4 – Atomic Models and Properties of Atoms

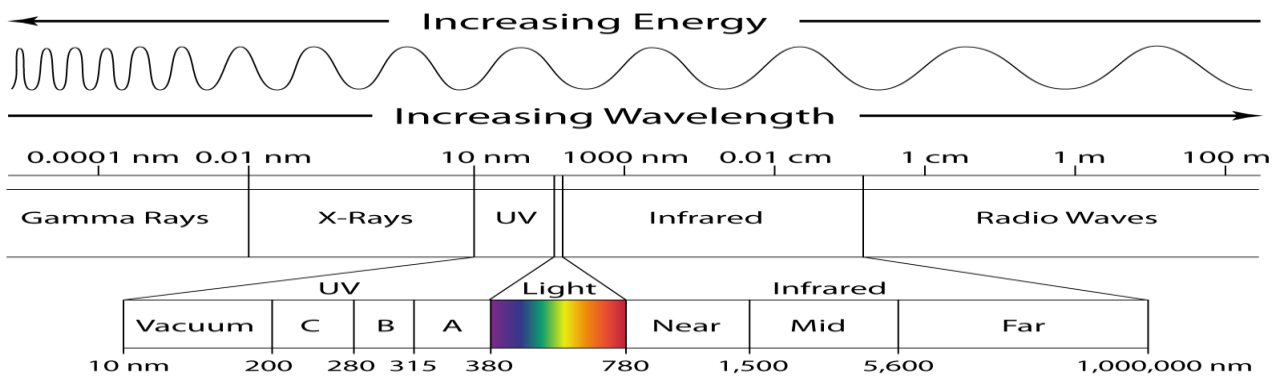
SCH 4UI Section 3.1 Nuclear Atomic Model and Emission Spectra Read p164-173 RQ #2-14

Review: Bohr-Rutherford Atomic Model indicates...

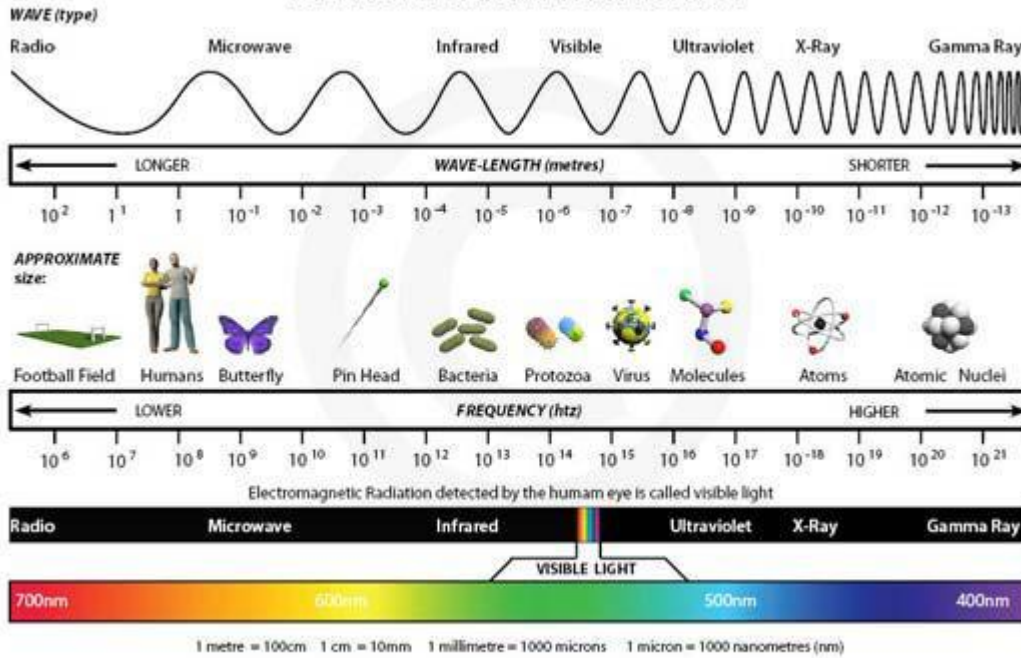
Example: Calcium

Bohr-Model page 171

- Electrons exist in circular orbits, much like planetary orbits, electrostatic force holds the e in orbit
- electrons can only exist in “allowed” orbits, only certain energy levels are allowed
- while an electron remains in one orbit, it does not radiate energy
- electrons can jump between energy levels by absorbing or emitting photons, these photons must have the exact amount of energy required to get to the higher level

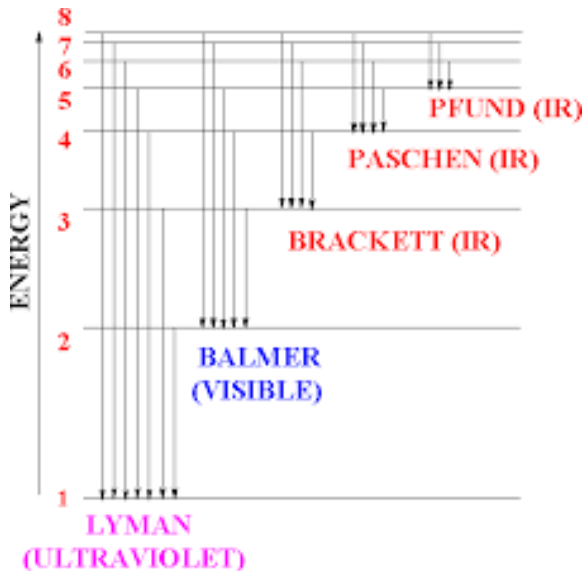
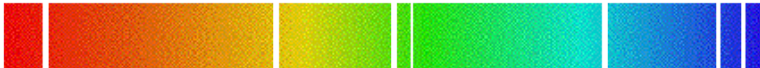
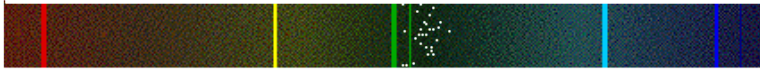


THE ELECTRO MAGNETIC SPECTRUM



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Emission and Absorption Spectrums ...



SCH 4UI Sect 3.2 Quantum Mechanical Model of the Atom

Hmwk Read Section 3.2 pages 174-179, pp#1-10

Prior Knowledge:

**Thompson's contribution, Rutherford and the Gold Foil Expt, Planck's Quantum of Energy, Bohr's Model, Emission Spectrum and Absorption Spectrum, Causes of UV, Visible light and Infrared*

The Basis of this model is to determine where the electrons are in an atom and what are they doing.

*Bohr's Model was too simple and could not explain all results, ie. why does Hg have 3 yellow lines very close together in it's emission spectrum, instead of 1 like that of H

New Idea was proposed by **Louis de Broglie, to treat matter; anything that has mass and volume as having wave like properties

*Note: Large objects with mass and volume the wave properties are negligible, but electrons are small therefore this theory has merit

Erwin Schrodinger combined de Broglie and Einstein's Ideas

***electrons have wave properties**

***Quantized Energy particles (photons)**

=Quantum Mechanical Model of the Atom (obtained using mathematical equations)

Werner Heisenberg (1927)

*proposed the uncertainty principle: states that you can not determine the location or velocity of an electron at the same point in time

*Therefore using statistics chemists were able to determine approximately where electrons are in an atom

*Heisenberg developed wave equations, which lead to the concept of orbitals

***Orbitals**: describe the region of probable location of electrons, an electron cloud of probability with a 90% degree of accuracy

*There are 4 types of orbitals that we are going to study, each has a different shape

*S orbital:

*P orbitals:

*D orbitals

*F orbitals:

*The location of electrons is described using quantum numbers. Quantum numbers essentially provide the co-ordinates for each electron in an atom.

***Principle Quantum number, n**

*describes the energy level

***Second Quantum number, l**

*describes the orbital shape quantum number

***Magnetic Quantum number, m_l**

*indicates the sub level of the orbital

***Fourth Quantum number, m_s**

*spin quantum number

Practice...

Review; Identify the energy level and the orbitals present for that energy level for the following

n = 1

n = 2

n = 3

Fourth Quantum Number, m_s , spin quantum number

-the two possible values are $+\frac{1}{2}$ and $-\frac{1}{2}$

- $+\frac{1}{2}$ represents a clockwise spin

- $-\frac{1}{2}$ represents an anticlockwise spin

-usually the + spin is written first

Pauli Exclusion Principle – an orbital may have a maximum of 2 electrons
and / OR No two electrons in an atom can have the same four quantum numbers

Example: Helium

Emission Spectrums, <http://astro.u-strasbg.fr/~koppen/discharge/>

*Single Electron System

Hydrogen



*Multi Electron System

Mercury



Remember that each element produces its own unique emission spectrum, **WHY?**

*What can you conclude about mercury's two yellow lines being very close together? What can you conclude about the energy levels or orbitals?

***One Electron System**

***Multi Electron System**

Aufbau Principle: is an imaginary process of building up the ground state electronic structure for each atom

Therefore use the following to aid in placing the electrons in the lowest orbital with respect to energy, this represents the ground state configuration.

Example:

Writing Electron Configurations...

-This is one short hand method of writing electron configuration that shows the n , l and the number of electrons for each orbital

Example:

***Condensed Form**, uses the previous noble gas

*Another method is using **Orbital Diagrams aka Orbital Box Diagrams**

-this represents a visual of the electron arrangement

-If $l = 0$, $m_l = \underline{\hspace{1cm}}$ therefore there would be $\underline{\hspace{1cm}}$ box

-If $l = 1$, $m_l = \underline{\hspace{1cm}}$, therefore there would be $\underline{\hspace{1cm}}$ boxes attached

-If $l = 2$, $m_l = \underline{\hspace{1cm}}$, therefore there would be $\underline{\hspace{1cm}}$ boxes attached

-If $l = 3$, $m_l = \underline{\hspace{1cm}}$, therefore there would be $\underline{\hspace{1cm}}$ boxes attached

Guidelines for "Filling" Orbitals, see page 184

1. Place electrons into the orbitals in order of increasing energy level
2. Each set of orbitals of the same energy level must be completely filled before proceeding to the next orbital or series of orbitals
3. Whenever electrons are added to orbitals of the same energy sublevel, each orbital receives one electron before any pairing occurs
4. When electrons are added singly to separate orbitals of the same energy, the electrons must all have the same spin

Hund's Rule: An extension of Pauli Exclusion Principle,

*No 2 electrons can have the same set of quantum numbers

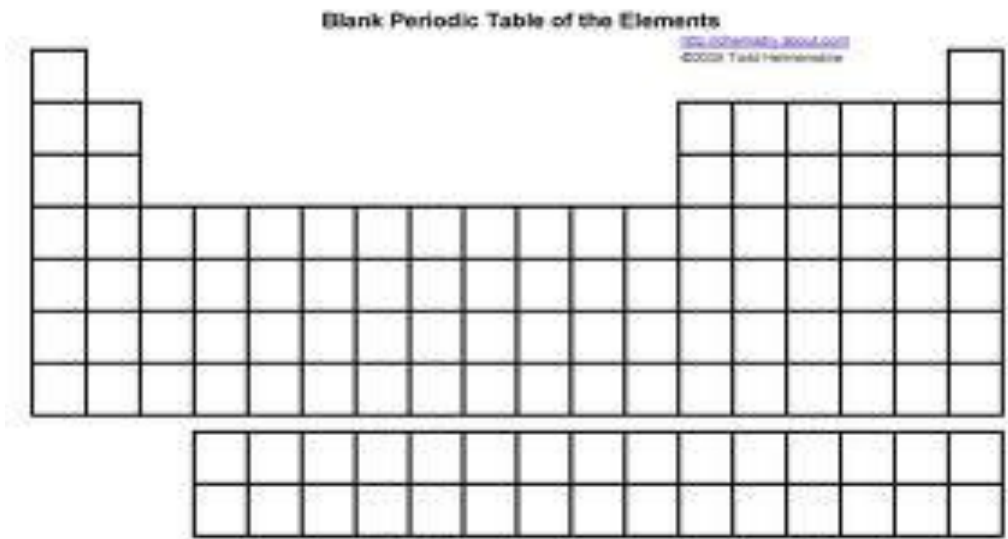
Example:

The **Periodic Table** is organized into 4 blocks...

Note:

- Valence electrons in the s and p orbitals form the main group elements ie metals, metalloids and non metals
- Valence electrons in the d-orbital for the ...
- Valence electrons in the f-orbital form the ...

Periodic Table:



*Elements in the same group have similar electron configuration and have similar chemical properties.

*When writing electron configuration the highest n value indicates the period the element is in

*Addition of the last s, p and d values will indicate the group number for the element

Other Useful Calculations...

n^2 = the number of orbitals in an energy level

$2n^2$ = the maximum number of electrons permitted in the energy level

Exceptions to Aufbau's Principle...

*Chromium

*Copper

-Draw the expected electron configuration for these 2 elements

- The real electron configuration as determined by experiment is... Why?

Trend 1 – Atomic Radius

-is measured by determining the distance between the nuclei of bonded neighbouring atoms

Remember the basic trend?

*Across a Period from Left to Right...

*Down a Group ...

WHY? Three factors that are need to be considered are...

1) **n, energy level** : As n increases the size of the orbitals increase and hence the electrons are further away from the nucleus

2) **Z and Zeff**

3) **Shielding**, the inner electrons shield the outer electrons from the effective nuclear charge the nucleus produces

Across a Period from L to R

*Down a Group

Trend 2 – Ionization Energy

-Is the amount of energy required to remove an electron from a ground state gaseous atom

1st IE, is the amount of energy required to remove the 1st electron, ie the valence electron

2nd IE, is the amount of energy required to remove a 2nd valence electron

Therefore the expected trend for Ionization Energy is...

*Draw the electron configurations for the following;

Mg and Al

N and O

Ne and Ar

*First ionization energy is closely linked to an atom's chemical reactivity, if an atom loses electrons easily, therefore has a low 1st ionization energy, than it would make sense that the atom would be highly reactive.

A good example of that are the elements in group 1.

***Chemical Reactivity** of Metals tends to increase down a group and decrease across a period from left to right

Trend 3 – Electron Affinity

-Is defined as the change in energy that occurs when an electron is added to a gaseous atom. These atoms would then be forced to form anions.

- The trends that arise from electron affinity vary and exceed the expectations of this course but

In general...

*Down a group Electron Affinity...

Summary:

Elements	Ionization Energy	Electron Affinity
Elements in Group 1 and 2		
Elements in Group 17		
Elements in Group 18		

Chemical Bonds are defined as “electrostatic forces that hold atoms together in compounds”.

Type 1 Ionic Bonds

- Formed due to the attraction of cations and anions
- often have large ΔEN
- the result of the ionic bonds are rigid lattice structures, arranged to have the lowest energy state possible
- Often the s orbital electrons of the metal are given to the p orbitals of the non metals

Formation of the ionic bond of sodium and chlorine

Properties of Ionic Solids:

- Crystalline solids at room temp
 - Brittle
 - Non conductors of electricity when in solid form
 - Can conduct electricity when dissolved in water and in liquid form
 - o WHY?
- Non conductors of heat
-High melting points
-Soluble in water except for MgF_2

Lattice Energy is the amount of energy released when ionic crystals form from the gaseous ion

*The compound with the largest lattice energy has the highest melting point

Type 2 Covalent Bonding

- Covalent bonds form as a result of electrons being shared between two atoms
- It is a balance between forces of attraction and repulsion that act between the nuclei and electrons of two or more atoms
- There is an optimum separation that favours a minimum energy for the system

HOW?

- When 2 half filled atomic orbitals from two atoms overlap to share the same region of space a covalent bond is formed
- These new orbitals are lower in energy and are therefore more energetically favourable

Examples:



Bond Energy is the energy required to break the force of attraction between two atoms in a bond
-If the number of bonds increases, the bond energy increases,

Properties of Covalent Molecules

- Solid, Liquid or Gas at room temp
- Low bpts and mpts
- Poor conductors of electricity and heat
- Poor solubility in water

Figure 4.2 and Figure 4.2 page 210 ... ΔEN

Type 3 Metallic Bonding

- Metallic bonds form due to a sharing of electrons between other metals
- Remember metals have low IE and low EA, as a result a “sea of electrons” or “Free electron model” forms

Properties of Metals

- Conductors of electricity and heat in solid and liquid state
- Malleable and ductile
- Change state at moderate to high temp

Type 2 Covalent Bonding Continued.... Advanced Concepts

-**Valence Bond Theory (VBT)** : atomic orbitals of 2 atoms overlap to form a covalent bond

-**Molecular Orbital Theory (MOT)**: orbitals overlap and combine to form new orbitals called molecular orbitals

-**Hybrid Orbitals**: are orbitals that are formed by the combination of 2 or more orbitals of valence shells of an atom

Ex H_2

Ex CH_4

Ex C_2H_4

Ex C_2H_2

SCH 4UI Section 4.2 Part 1 Lewis Structures Plus *Hmwk Read pages 228-232, pp 1-10*

*Lewis structures are used to show how molecules and or polyatomic ions are bonded together, ie single, double or triple bonds. Lewis structures also identify the number of lone pairs of electrons in the molecule... this is important in the next lesson...

How to draw **Lewis Structures, ie Rules / Guidelines**

1. The least electron negative atom is in the center; hence the higher electronegative atoms surrounds the least EN atom
-Note: H is always terminal, remember can only form one single bond
2. Count the total number of valence electrons for each atom
-If the molecule has a charge (polyatomic ions) the charge must be incorporated into this count
-If the charge is negative, add the charge to the number of valence electrons
-If the charge is positive, subtract the charge from the number of valence electrons
3. Determine the total number of electrons for each atom to obtain its octet
-Note: H needs only 2
-Subtract the number of valence electrons from the total number of electrons required, this value represents the number of shared electrons
-Divide the number of shared electrons b 2 to determine the number of bonds in the molecule ... this should help you to determine if the molecule has single, double or triple bonds
4. To determine the number of lone pair of electrons, the ones not involved in a bond, subtract the number of shared electrons from the number of valence electrons. This is the number of electrons that are available to be placed around atoms to obtain their octet

Good Luck counting electrons and figuring this all out ☺

Practice:



Resonance Structures, are structures or molecules with more than one possible arrangement of bonding

Co-ordinate Covalent Bonds, occur when one atom contributes both of its electrons to form a covalent bond



Expanded Octets, aka Expanded Valence Levels, this occurs when the central atom has more than 8 electrons surrounding it.

AND Incomplete Octets:

SCH 4UI Sect 4.2 con't ... VSEPR – Valence Shell Electron Pair Repulsion Theory

Reference 2002 text pages 178-186, 2011 text pages 232-236 pp#11-20

-This theory was created by Canadian Chemists at McMaster University, Ronald Gillespie and Ronald Nyholm

-this model is based on the fact that pairs of electrons that surround the central atom in a molecule repel each other and arrange themselves in space in such a way that they are as far apart as possible

Key Point

-Lone pair of electrons take up more space than bonded pair of electrons

Therefore;

LP-LP > LP-BP > BP-BP

.... This explains why the angles between atoms within a molecule change if there are lone pairs of electrons on the central atom

Examples:

-methane

ammonia

water

Definitions for VSEPR table:

***Geometric arrangement of electron groups**, refers to the groups of valence electrons and how the valence electrons are positioned

***Molecular Shape**, the shape is based on the position of the actual atoms themselves in the molecule and ignores the position of the non bonded electrons

VSEPR Notation:

A – represents the central atom

X – represents the bonded pair of electrons

E – represents the lone pair of electrons that surround the central atom

Note: double and triple bonds count as one electron group

Summary of VSEPR...

# of e ⁻ groups	Geometric arrangement of e ⁻ groups	# lone e ⁻ / # of bonded e ⁻	VSEPR notation	Molecular shape	Bond angle(s)	Sketch

Practice:

1. Use VSEPR theory to predict the molecular shape for each of the following;

- a) HCN b) SO₂ c) SO₃ d) SO₄²⁻ e) AsCl₅ f) BF₄⁻
g) XeI₂ h) PF₆⁻ i) AsF₃ j) AlF₄⁻

2. Given the general formula and the shape of the molecule or ion, suggest possible elements that could be the central atom, X, in each of the following;

- a) XF₃⁺ (trigonal pyramidal) b) XF₄⁺ (tetrahedral) c) XF₃ (T-shaped)

SCH 4UI Section 4.2 Part 3 Intermolecular Forces

Hmwk Rd 238-247 Review Questions #7, 11-14

Intramolecular Forces, are the forces within a molecule or polyatomic ion

-during a chemical change the intramolecular forces are overcome and chemical bonds are broken and others are formed

-these forces influence chemical properties of substances

Intermolecular Forces are forces of attraction and repulsion that occur between molecules or ions

-these forces influence physical properties of substances

Therefore, intermolecular forces tend to be weaker than intramolecular forces.

Remember ... ΔEN ... represents the type of bond, ie intramolecular forces

POLAR molecules OR NON POLAR molecules, How do you know?

-Polar or Polarity, means an uneven distribution of charge, resulting in polar bonds and maybe polar molecules

-Polar bonds occur when the difference in electronegativity between two atoms falls in the region of 0.5 to 1.7

Use the chart below to help you decide if the molecule is Polar OR Non polar overall...

POLAR Molecules

-Must have polar bonds

-Forces on the bonds do NOT cancel

*This occurs when...

-different types of atoms surround the central atom

-there are lone pairs of electrons on the central atom ... usually

NON POLAR Molecules

-do NOT have polar bonds within

-Forces cancel each other out

-same type of atom surrounds the central atom the force will cancel out

Other types of Intermolecular Forces:

1) Dipole – Dipole

2) Ion – Dipole Forces

3) Ion induced dipole forces

4) Dipole induced dipole forces

5) Dispersion Forces

6) Hydrogen Bonding

Bonding in Solids...Text 2002 see **pages 196-207**, Text 2011 see **pages224-225**

*Solids can be classified according to molecular or particle arrangement:

1) Crystalline Solids

*organized arrangement

*distinct shapes

2) Amorphous Solids

*indistinct shapes

*random order

**5 types of Crystalline Solids: Atomic, Molecular, Network, Ionic and Metallic

Metallic...see notes

Atomic Solids:

*These are solids made up of individual atoms, therefore the only intermolecular force holding these solids together are dispersion forces. The only example of atomic solids are noble gases at low temperatures, which have a filled octet. Hence these molecules experience low melting points and boiling points.

Molecular Solids:

These are solids made up of molecules...perhaps the one you have talked about the most so far in chemistry. This category includes non polar molecules making up non polar solids and polar molecules making up polar solids. Therefore these solids experience a wide range of intermolecular forces including dispersion, dipole-dipole and hydrogen bonding. Depending on the type of molecule and hence the type of intermolecular forces boiling points and melting points vary.

Network Solids:

These are atoms that "bond covalently into two or three dimensional arrays", meaning these atoms form massive molecules forming sometimes 3-dimensional shapes.

*A prime example is carbon...it can bond to form 4 very different structures.

Example 1...Graphite

*Carbon forms layers of connecting carbon atoms by forming 3 covalent molecules in a trigonal planar shape. As a result the "4th pair of electrons are not involved in the covalent bonds and are free to move. This is what enables graphite to conduct electricity. Another property of graphite is that it is somewhat soft and slippery. The layers of graphite are not held together by covalent bonds but are held together by dispersion forces...hence the graphite layers can slide across each other somewhat easily. An application of graphite is that it can be used as a lubricant for machinery.

Example 2...Diamond

*The arrangement of the covalent bonds with this carbon based network is the traditional tetrahedral arrangement, however the tetrahedral arrangement continues repeatedly forming a 3-dimensional structure. In this arrangement all electrons are held in position via the covalent bonds and therefore is unable to conduct electricity. The planes of the atoms are what causes light to be reflected, producing diamonds characteristic sparkle. A third characteristic of the diamond is the strength of the material, again this reinforces the incredible strength of the repeated covalent bond.

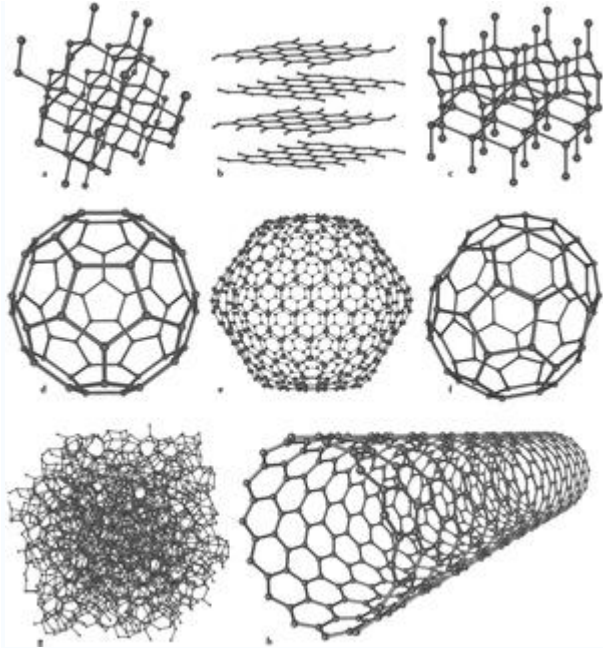
Example 3...Fullerenes

*Fullerenes were invented in 1985 and are made with 60 carbon atoms or more. Again these networks possess the tetrahedral arrangement of the diamond, but they do so to create a sphere. Infact they represent the same shape that was created by an architect, Buckminster Fuller...hence the name buckyballs. It is speculated that these network solids make good lubricants.

Example 4...Nanotubes

*Nanotubes are an extension of the fullerenes. Instead of a sphere being created the carbons arrange themselves forming tube.

Example 1-4 represent the term allotropes...molecules made from the same element but differ in arrangement.



Eight allotropes of carbon: diamond, graphite, lonsdaleite, C60, C540, C70, amorphous carbon and a carbon nanotube.

Other Network Solids...

* SiO_2 is the simplest formula that is repeated in a network solid, an example is Quartz.

Ionic Crystals:

The ions involved in forming ionic crystals arrange themselves in a regular pattern or grouping called a crystal lattice. The basic repeating pattern is called a unit cell. The size of the atoms dictate the type of unit cell that forms, ie. NaCl forms a face-centered cubic structure, where as CsCl forms a body centered cubic structure.

