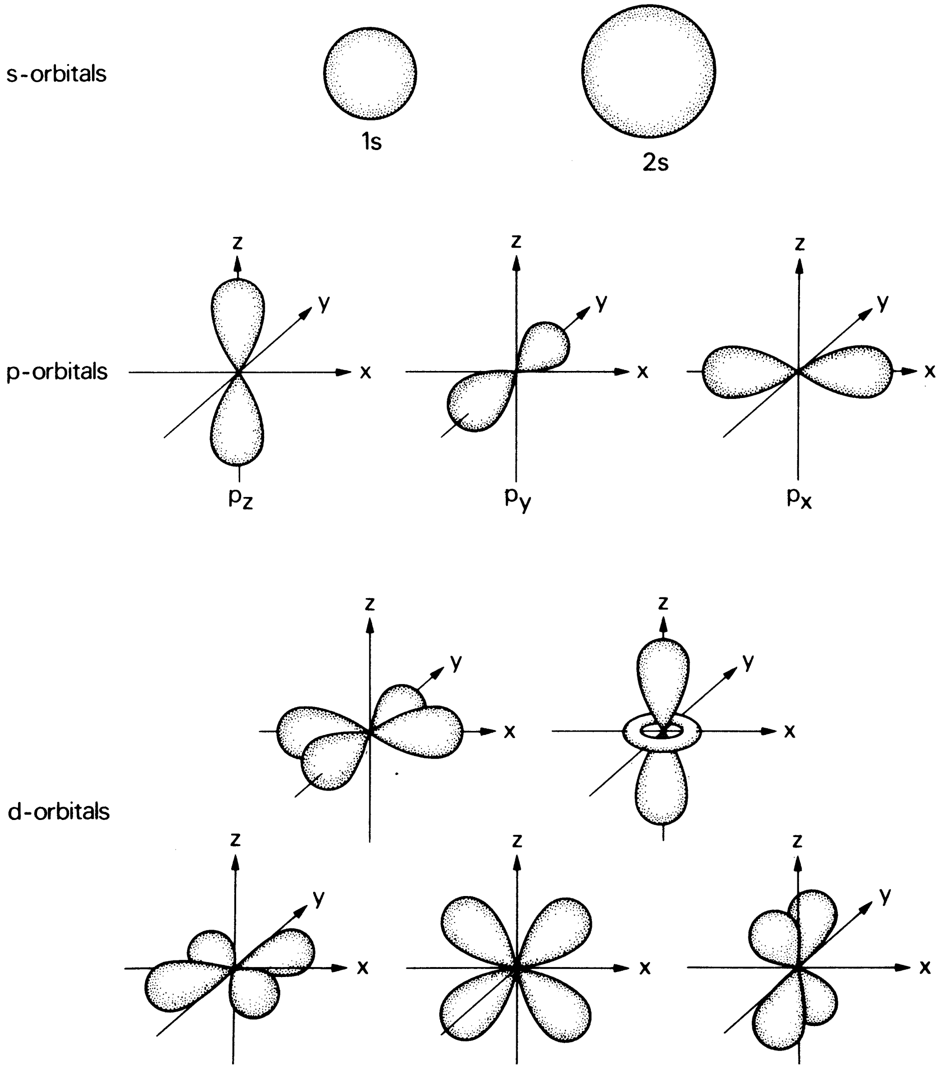
SCH 4U1

UNIT 1

STRUCTURE AND PROPERTIES OF MATTER

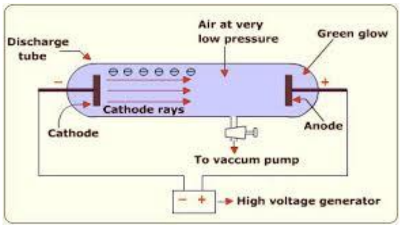


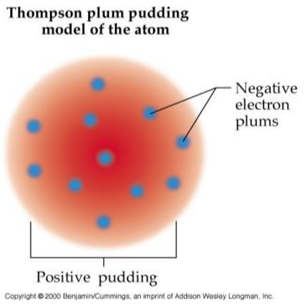
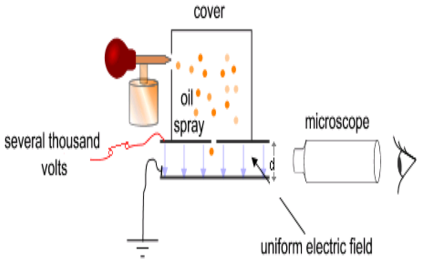
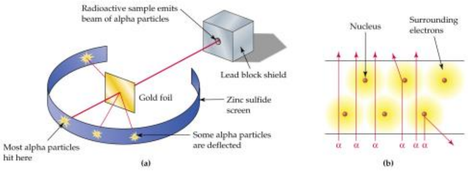
SCH 4U – UNIT 1: Structure and Properties Matter Timeline

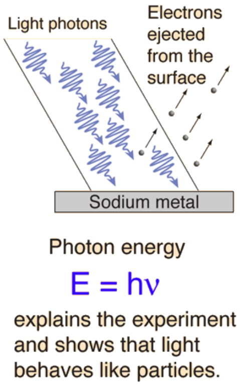
DAY	TOPIC	HOMEWORK
	INTRODUCTION Review of Grade 11 Chemistry/Lab Safety/Lab Equipment	Review Package
	TOPIC 1: Development of the Atomic Theory Bohr's Atomic Theory and Development of Quantum Mechanics	**Read 3.1** Do p. 142 # 3-7 **Read 3.2** Do p. 147 # 4-8
1	**TOPIC 2AB**: Quantum Mechanics and Quantum Numbers Shapes of Orbitals	**Read 3.3 & 3.4** Do p. 152 1 & 3; p. 158 # 1, 2 p. 159 # 4, 6, 7, 8, 9 & 11 **Extra Practice Worksheets** **Quantum Numbers Worksheet.pdf** **Quantum Numbers Worksheet key.pdf**
1	**TOPIC 3:** Representing Electrons: Energy Level Diagrams/Orbital Block & Electron Configuration Applications of Quantum Mechanics	**Read 3.5** Do p. 166 # 1 – 4; p. 170 # 1- 2; p. 172 # 1c,d, 4,5, p. 176 # 5,6 **Extra Practice Worksheets** **Energy Level Diagrams and Electron Configuration** **Energy Level Diagrams and Electron Configuration key.pdf** **Orbital Block Diagrams and Electron Configurations** **Orbital Block Diagrams and Electron Configurations key.pdf** **Electron Configuration Practice Worksheet and Answer Key.pdf**
2	**TOPIC 4**: Review – Trends, Chemical Bonding **TOPIC 5:** Lewis Dot Diagrams, Calculating Formal Charges, Resonance Structures **TOPIC 6:** VSEPR theory – shapes of molecules and predicting shapes Polarity of Molecules using Shapes	**Read 4.2** Do p. 200 #1-2, p. 204 # 1-2, 212 # 1, 2; p. 214 #1; p. 215 # 1, 2; p. 216 # 2-5,8 **Read 4.5** Do p. 227 #1-3, p. 229 #5-7 **Read 4.3** Do p. 220 # 1,2 p. 221 # 1, 6, 10, **VSEPR SHAPES ANIMATION** **Lewis Structures and Formal Charges Practice**
3	**TOPIC 6B Cont'd:** Hybrid Orbitals – Using energy level diagrams Single, double and triple bonds **TOPIC 7:** Coordinate Covalent bonding	**Read 4.6** 238 # 2,3, 5, 6, 8 & 11 **Identify Hybridization Mini quiz** **Extra Practice Worksheets** **Hybridization Worksheet.pdf** **Hybridization Worksheet key.pdf**

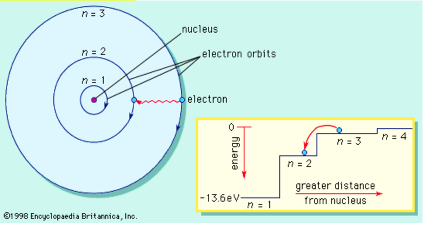
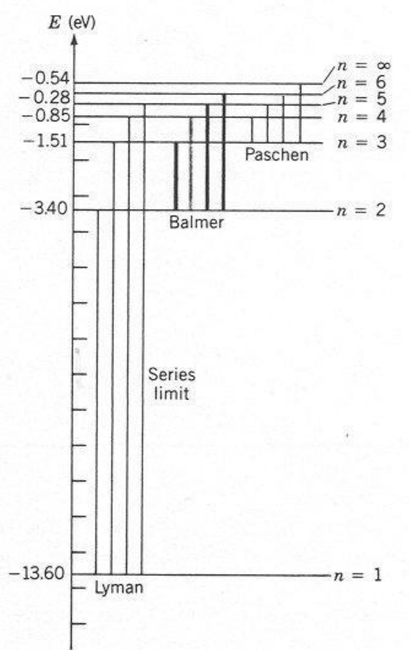
		🎥 How to Identify Hybridization
3	**TOPIC 8**:Intermolecular forces – LD, DD, hydrogen bonding, ionic, network, metallic	**Read 4.7** Do p. 244 # 1,2; p. 247 # 1(not c), 2 (not c/d)-8 **Extra Practice Worksheets** **Chemical Bonding Worksheets.pdf** **Chemical Bonding Worksheets Key.pdf**
4	**TOPIC 9**: Aggregates – their properties Ionic crystals, metallic crystals, molecular crystals, network crystals	**Read 4.8** Do p. 254 # 2 – 11 **Properties of Solids mini quiz** **Extra Practice Worksheets** **Structure, Bonding, and Properties Worksheets.docx** **Structure, Bonding and Properties Key.pdf** 🎥 Identifying Types of Solids Overview
5	**Test**	**Review**: Review at the back of package notes **Structures & Properties Review Key** Any of the questions from text pp 272 – 278 # 1 - 123; **Plus** p 278 # 126, 127, 128

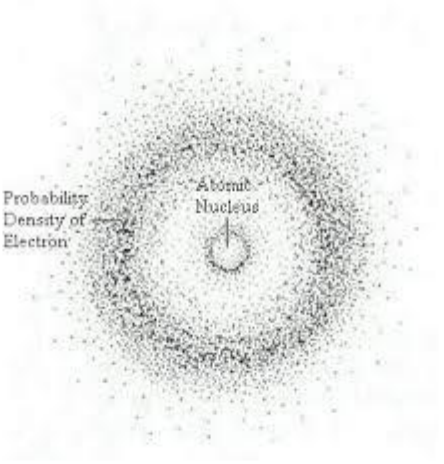
TOPIC 1: THE DEVELOPMENT OF THE ATOMIC THEORY

Scientist	Contribution
History of an Atom Democritus (300 B.C.)	• Hypothesized that matter cut into smaller and smaller pieces would eventually reach the atom - idea disbelieved/ignored for hundreds of years
Robert Boyle	• re-introduced Democritus' idea about matter
John Dalton (1805)	• laws of definite composition (specific combining capacity), multiple proportions (some atoms have more than one combining capacity) and conservation of mass (matter neither created nor destroyed in a chemical reaction) • Matter is composed of indestructible, indivisible atoms which are identical for one element but different from other elements
Svante Arrehnius	• atoms may gain or lose electrons to form ions in solution
Michael Faraday (1832)	• particular atoms and ions gain or lose a specific number of electrons • coined the term "electrolysis" as a splitting of molecules with electricity
Dmitri Mendeleev (1869)	• arranged elements into 7 groups with similar properties • discovered that the properties of elements "were periodic functions of their atomic weights" – known as the Periodic Law
William Crookes (1879) 	• invested sealed glass tube with 2 electrodes • when tube filled with a gas, and partially evacuated, electricity would flow from cathode (+ terminal) to anode (- terminal) • colour of glow dependent on type of gas used • when tube completely evacuated, green ray still visible • suggested that electricity is composed of negatively charged particles

J. J. Thomson (1887)  	• electrons are a component of all matter – hypothesized that the atom was composed of electrons embedded in a positively charged sphere (raisin bun model of atom)
E. Goldstein (1886)	• discovered a ray which went in the opposite direction to the cathode ray when in a magnetic or electrical field applied • suggested that another subatomic particle existed
W. Roentgen(1895)	• while working with discharge tubes discovered X-rays
H. Becquerel (1896)	• discovered that some materials gave off invisible rays – called these materials radioactive
Robert Milliken (1909)  	• determined the charge on an electron using the oil drop experiment (sprayed a mist of oil drops into a chamber between 2 electrical plates; drops became charged and the charged droplets were suspended between the plates by balancing the electrical force upward from the plates with the downward gravitation force • determined the charge of an electron to be 1.602 X 10^{-19} coulombs and a mass of 9.11 X 10^{-28} g
Ernest Rutherford (1898)  	• worked with alpha, beta and gamma radiation used radium as a source of alpha radiation and directed it onto a 300-atom thick piece of gold foil • predicted based on Thomson's model that the alpha particles would go right through the foil, some did but some reflected and deflected • proposed the nuclear model of the atom i.e. Atom consists of a very dense positively charged centre called the nucleus – this takes up very little volume while the electrons orbit around the nucleus and make up most of the volume of the atom • proposed that the atom's nucleus was held together by nuclear attractive forces • in 1914 studies resulted in the identification of the proton

Francis Aston (1919)	• developed the mass spectrometer to identify isotopes
Fredrick Soddy (1913)	• identified the existence of isotopes, discovered "half-life" and performed initial decay calculations • proposed that isotopes of an element had the same number of protons
Bunsen and Kirchhoff (1859)	• developed the spectroscope • studied the light emitted by blackbodies (bodies that do not reflect any light) and light given off by heated objects • studied emission and absorption line spectra for gaseous elements • elements burn with a characteristic colour and give a characteristic bright line spectrum can identify different elements • absorption spectrum – dark-line spectrum places a gas between the continuous spectrum source and the observer with the result that the gas's dark lines appeared in the same place as where the lines would be in a bright line spectrum; emission spectrum – series of coloured lines of light produced or emitted by a gas excited by heat or electricity
Max Planck (1900)	• started the **quantum revolution** • developed the mathematical equation to explain the curve produced when an objected is heated • hypothesized that the energies of the oscillating atoms in the heated solid were multiples of a small quantity of energy (or quantum) – i.e. that energy is not continuous but is in bursts where the little packet of energy is a quantum of energy
James Maxwell (1873)	• theorized that light is an electromagnetic wave composed of electric and magnetic fields that can exert forces on charged particles – this is the electromagnetic-wave theory (classical theory of light) i.e. light is an electromagnetic wave consisting of a continuous series of wavelengths
Heinrich Hertz 	• accidentally discovered the photoelectric effect in 1887 • photoelectric effect is the effect that electromagnetic radiation (light) has on substances in particular certain metals • showed that the brightness (intensity) of light did not determine the kinetic energy of the liberated electrons from the metal but that the frequency (colour/energy) of light determined the kinetic energy of the liberated electrons from the metal

Albert Einstein (1905)	• hypothesized that light consists of packets of energy or quanta which he called photons • each colour of light has a specific amount of energy i.e. photon of one colour of light may have more or less energy than that of another colour of light • theorized that photoelectric effect is the result of a photon of light striking the metal's electron and transferring its energy to the electron causing the electron to leave the metal
Neils Bohr  ©1998 Encyclopaedia Britannica, Inc.  **The Bohr Model**	• had to solve the problem of electrons emitting photons of electromagnetic radiation (accelerating charges continuously produce some type of light) thus losing energy and so spiralling in toward the nucleus and collapsing • reasoned that if the light released or absorbed from an atom was **quantized**, then the energy of the electron inside the atom must be quantized –i.e. electrons can only have certain energies and that at ground state the energy is constant – **1st postulate** – electrons do not radiate energy as they orbit the nucleus – each orbit corresponds to a state of constant energy called the stationary or ground state • from the mathematical equation of Jacob Balmer, Bohr stated that electrons jump from one energy level to another (an energy transition) and so when an electron drops from a higher energy state to a lower energy state, there is a loss of energy which is seen as light (bright line spectrum) when an electron absorbs energy to move from a lower energy state to a higher, then the dark line spectrum is seen **–2nd postulate** – electrons can change their energy only by undergoing a transition ("jump") from one ground state to another; when an electron moves from a higher energy state to a lower energy state the energy is given off as light – **emission spectra**; when an electron moves from a lower to a higher energy state, it must absorb a certain amount of energy – **absorption spectra** (Photoelectric Effect) • using Planck's equation, Bohr was able to explain the experiment observations of hydrogen's spectrum • Bohr's model is able to explain Mendeleev's periodic table with periods being the result of energy levels • Proposed that an atom consists of electrons moving around a nucleus at high speeds in closed orbits (paths) at a definite distance form the nucleus; atoms contain a tiny central material called the nucleus where all the positive charges and most of the mass is located; a neutral atom has #p = #e; electrons are held in their orbits by the electrostatic force of attraction between the nucleus and the electrons • There are an infinite number of orbits available to the electrons

Louis de Broglie (1923)	• discovered the **duality of the electron** – ie it behaves like a particle and like a wave of energy
Werner Heisenberg (1927)	• rationalized that it was impossible to predict the exact location or velocity of atomic particles since they are so small and travel at the speed of light – called this the **"uncertainty principle"**
Erwin Schrödinger (1924)  	• devised **Wave Mechanics** (quantum mechanics) to account for the wave nature of the atom • the **mathematical expression** describes the **energy** and **motion of an electron** around a nucleus is called a **wave equation** • this expression gives the **probability of electron** being in a **certain place at a certain time** – the volume of space around the nucleus where the electron is most likely to be found is called an **ORBITAL** (electron cloud i.e. probability pattern) • solving the equation results in four quantum numbers • $\left(e^{-i\varphi}\psi(r_1,r_2)^*\right)\left(e^{i\varphi}\psi(r_1,r_2)\right)=\psi(r_1,r_2)^*\psi(r_1,r_2)$
James Chadwick (1932)	• identified the neutron which was also found in the nucleus

KNOW!!!

TOPIC 2A: QUANTUM MECHANICS (Q.M.)

Quantum Theory of an Atom and Quantum Numbers

- this model is purely mathematical and gives a probability of where the electron might be

The model:

a) was developed from the inability of measuring both the position and velocity of a body as small as electrons at the same time i.e. classical mechanics cannot be applied to an electron (**Heisenberg's Uncertainty Principle**)

b) now considers **the electron to be equal to a photon**. It has a **dual nature** – both particle- and wave-like **(de Broglie)**

c) replaces orbits with **ORBITALS** – a region around the nucleus where an electron will **probably be found** – the specific orbital that an electron will be in depends upon the **quantized energy of the electron** **NOTE** – the orbital **does not tell us where the electron is at any given time or how it moves in the orbital (Schrodinger)**

The mathematical equation describing the **energy and motion of an electron** around the nucleus is called the **WAVE EQUATION.** Solving the equation depends on these **four quantum numbers.**

TOPIC 2B: QUANTUM NUMBERS

Quantum Numbers

a) Principle Quantum Number, *n*

- identifies the energy possessed by the electron in any orbital and the distance from the nucleus – each orbital can **hold 2 electrons**
- **n = 1 - ∞**
- **for every value of n, there are n types of orbitals(sublevels) and n^2 actual orbitals**

e.g. **When n = 1**

- there is **n = 1**, one type of orbital and $n^2 = (1)^2 = 1$ actual orbital possible
- this type of orbital is called the **"s"** orbital
- the **shape** of the electron cloud produced by the moving electron is **spherical** (i.e. the electron is somewhere in the spherical shaped orbital at any given time because it possesses a specific amount of energy)
- the **s orbital** is found in all energy levels, each s orbital can hold a **maximum of 2 electrons**

When n = 2

- there are n**=2**, two types of orbitals and $n^2 = (2)^2 = 4$ actual orbitals possible in this energy level
- one of these orbital types is the **s orbital**, the other three are the **p orbitals**
- electrons with this energy generate electron clouds that have a **dumb-bell shape**
- the p orbitals exist in **every energy** level **from 2 on up**
- there are **three different orientations that p orbitals have** – these are relative to the axis on which they lie **(p_x, p_y, p_z)**
- since each orbital can hold a maximum of 2 electrons, **the p orbitals hold a total of 6 electrons**

When n = 3

- there are **n=3**, three types of orbitals and $n^2 = (3)^2 = 9$ actual orbitals possible in this energy level
- one of these is the **s orbital**, three are the **p orbitals**, and the remaining **5 are called the d orbitals**
- the electron clouds of the **d orbitals are butterfly shaped**
- the d orbitals exist in **every energy** level from **3 on up**
- the **five different d orbitals** are located in planes around the nucleus
- since each orbital can hold a maximum of 2 electrons**, the d orbitals hold a total of 10 electrons**

When n = 4

- there are **n=4,** four types of orbitals and $n^2 = (4)^2 = 16$ actual orbitals possible in this energy level
- one of these is the **s orbital**, **three are the p orbitals**, **5 are the d orbitals** and **seven are the f orbitals**
- the electron clouds of the **f orbitals are rosette shaped**
- the f orbitals exist in **every energy** level from **4 on up**
- since each orbital can hold a maximum of 2 electrons, **the f orbitals hold a total of 14 electrons**
- the **next type of orbital is the g orbital**, according to Q.M**. there would be 9 different g orbitals holding a maximum of 18 electrons**

NOTE: the maximum # of electrons in any given **energy level (n)** is given by **$2n^2$**

b) Secondary Quantum Number (Azimuthal/Angular Momentum), ℓ

- **where ℓ = 0, 1, 2, 3 ... n - 1**
- identifies the **SHAPE of the orbital**
- **when ℓ = 0; s orbital**
- **when ℓ = 1; p orbitals**
- **when ℓ = 2; d orbitals**
- **when ℓ = 3; f orbitals**
- **when ℓ = 4; g orbitals etc**.
- e.g. What are the possible values for ℓ when n = 3?

ℓ= 0
ℓ= 1
ℓ=0…. (n-1) ℓ=2

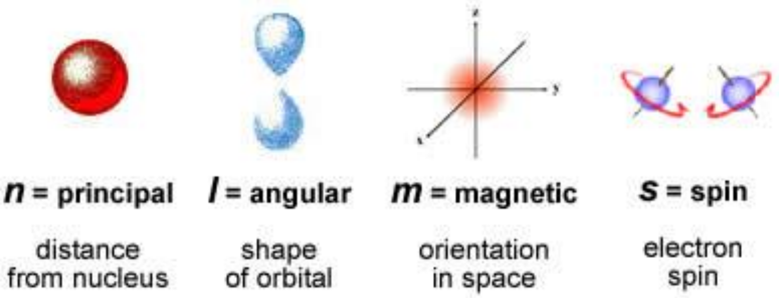
- as value of n increases, the size of the orbital also increases i.e. 1s is smaller than 2s which is smaller than 3s etc.

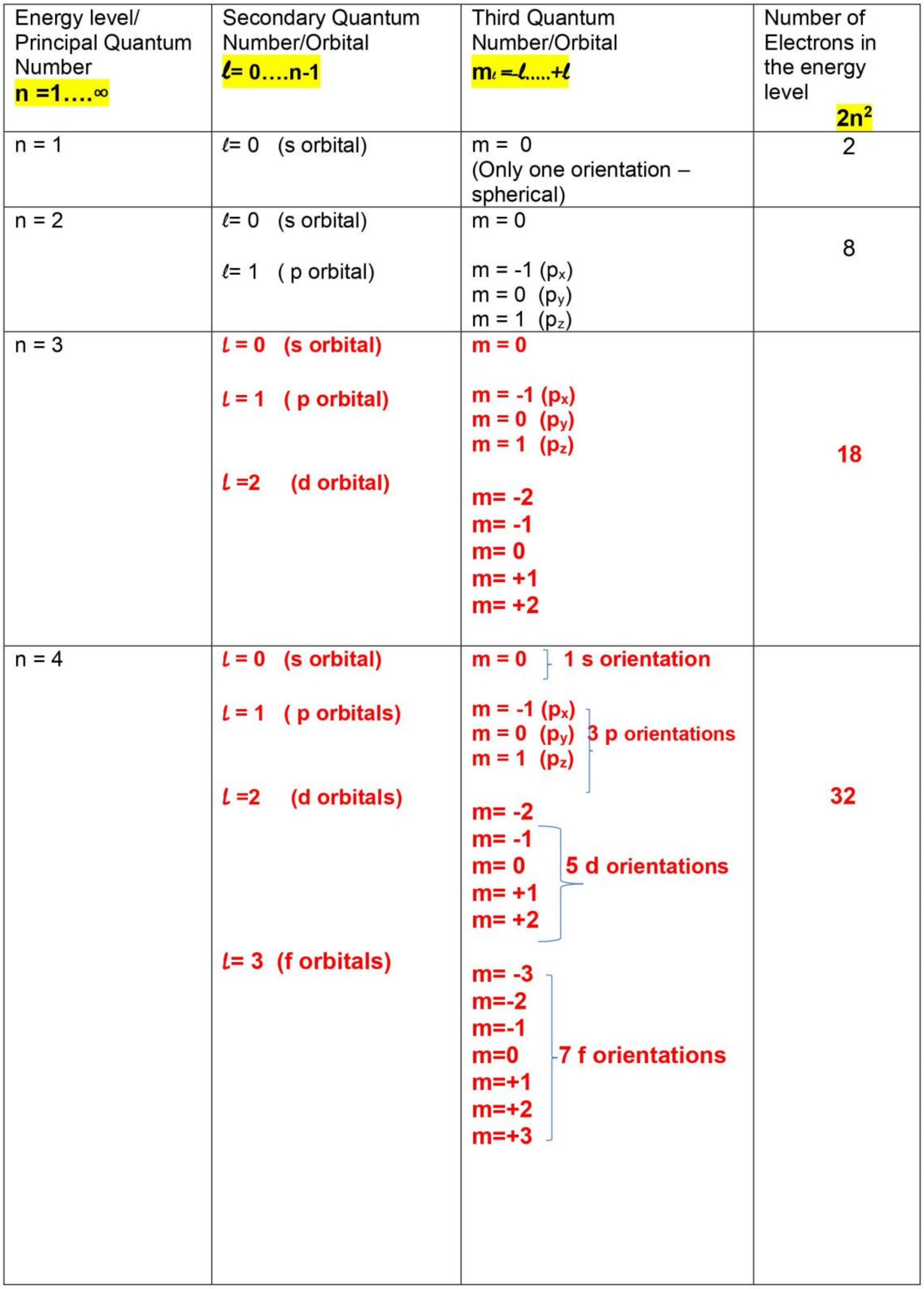
c) Third Quantum Number (Magnetic), *m or m_l*

- identifies the direction of orientation of the orbital with respect to external magnetic field
- for **any value of l,** m takes every integral value starting at **- l to + l**
- eg. l= 0; m = 0
  l= 1; m = -1; m= 0; m = +1

d) Fourth Quantum Number (Electron Spin), *s or m_s*

- has **two values + ½ and - ½** – indicates the direction of electron spin – the electron can spin either **clockwise** or **counter clockwise**
- if two electrons are in the same orbital, **they must have opposite spin**





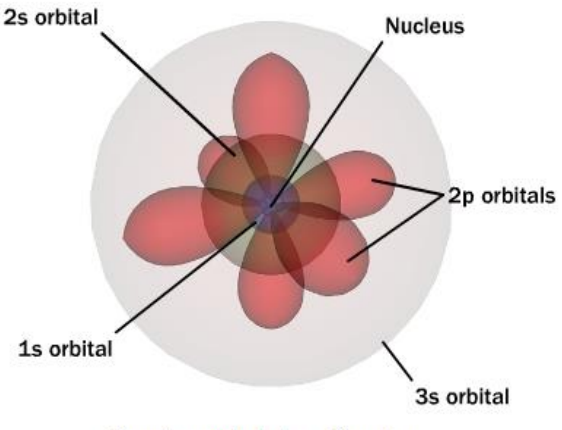
Energy level/ Principal Quantum Number **$n = 1 \ldots \infty$**	Secondary Quantum Number/Orbital **$\ell = 0 \ldots n-1$**	Third Quantum Number/Orbital **$m_\ell = -\ell \ldots +\ell$**	Number of Electrons in the energy level **$2n^2$**
n = 1	$\ell = 0$ (s orbital)	m = 0 (Only one orientation – spherical)	2
n = 2	$\ell = 0$ (s orbital) $\ell = 1$ (p orbital)	m = 0 m = -1 (p_x) m = 0 (p_y) m = 1 (p_z)	8
n = 3	**$\ell = 0$ (s orbital)** **$\ell = 1$ (p orbital)** **$\ell = 2$ (d orbital)**	**m = 0** **m = -1 (p_x)** **m = 0 (p_y)** **m = 1 (p_z)** **m= -2** **m= -1** **m= 0** **m= +1** **m= +2**	**18**
n = 4	**$\ell = 0$ (s orbital)** **$\ell = 1$ (p orbitals)** **$\ell = 2$ (d orbitals)** **$\ell = 3$ (f orbitals)**	**m = 0** } **1 s orientation** **m = -1 (p_x)** **m = 0 (p_y)** } **3 p orientations** **m = 1 (p_z)** **m= -2** **m= -1** **m= 0** } **5 d orientations** **m= +1** **m= +2** **m= -3** **m=-2** **m=-1** **m=0** } **7 f orientations** **m=+1** **m=+2** **m=+3**	**32**

COMPARISON OF BOHR AND Q.M. MODELS

KNOW!!!

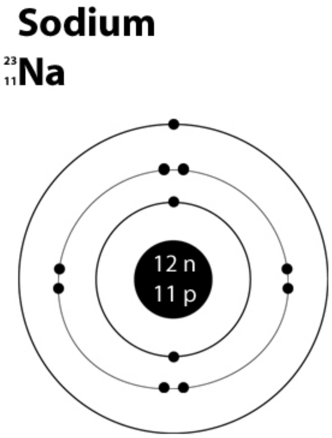
	BOHR	QUANTUM MECHANICS
nucleus	Same (p^+ and n^0)	Same (p^+ and n^0)
dimensions	2 - dimensional	3 - dimensional
electron characteristics	particle	wave and particle
electron motion	distinct paths (orbits)	probability of location (orbitals)
sub-levels	no	yes, s, p, d, f, etc
electrons	$2n^2$	two electrons per orbital so that $2n^2$/energy level

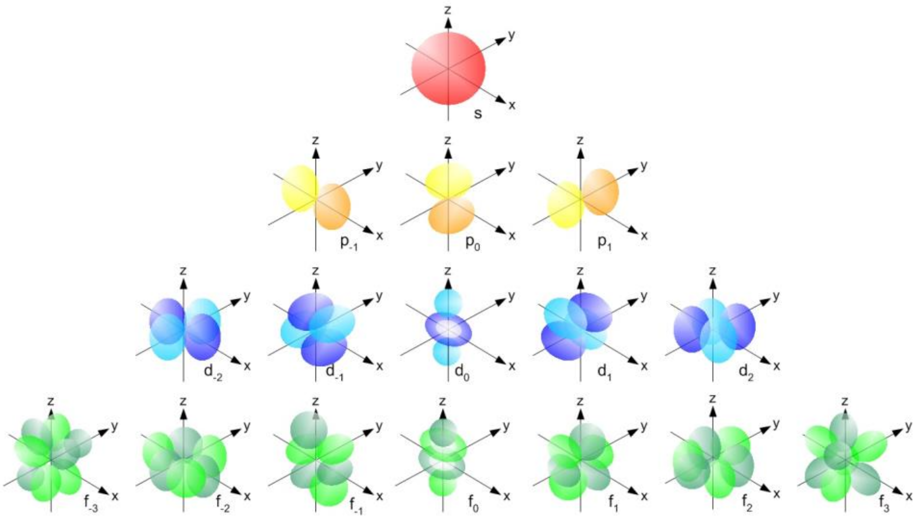
ORBITAL REPRESENTATION OF SODIUM



Quantum model of a sodium atom.

BOHR- RUTHERFORD REPRESENTATION OF SODIUM





TOPIC 3A: ORBITAL ENERGIES AND THE PERIODIC TABLE

- orbitals in an atom overlap
- **the orbitals have different amounts of energy**
- the energy from **lowest** to **highest** is:

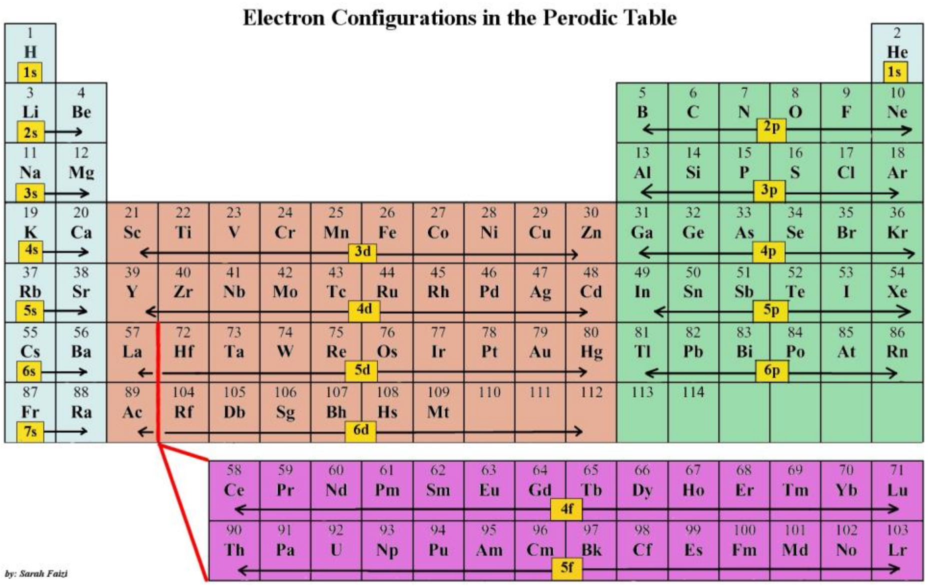
1s < 2s ⩽ 2p < 3s < 3p < 4s ⩽ 3d < 4p < 5s ⩽ 4d < 5p < 6s ⩽ 4f ⩽ 5d < 6p <7s < 5f < 6d < 7p

lowest energy → **higher energy**

Two ways to remember this –

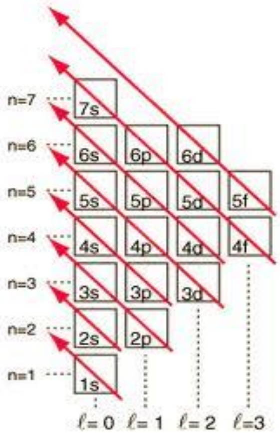
USE THE *PERIODIC TABLE

Orbitals on the Periodic Table



Memorize this !

OR use **a mnemonic**



TOPIC 3B: REPRESENTING ELECTRON ARRANGEMENT

Three rules control the filling of orbitals with electrons: KNOW!

The Electron Filling Analogy

1. **Aufbau Principle** – electrons enter orbitals of the lowest energy level first
2. **Pauli Exclusion Principle** - an atomic orbital may hold a maximum of 2 electrons, for two electrons to be in the same orbital, they must have opposite spin
3. **Hund's Rule** – when electrons occupy orbitals of equal energy, one electron enters each orbital until all the orbitals contain one electron with parallel spin. The second electron then adds to each orbital so that their spin are paired and opposite in direction

Different electron representations exist:

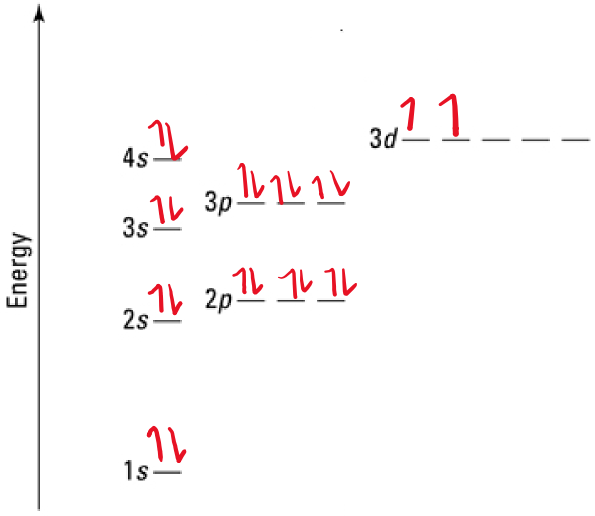
a) *****Energy Level Diagrams**

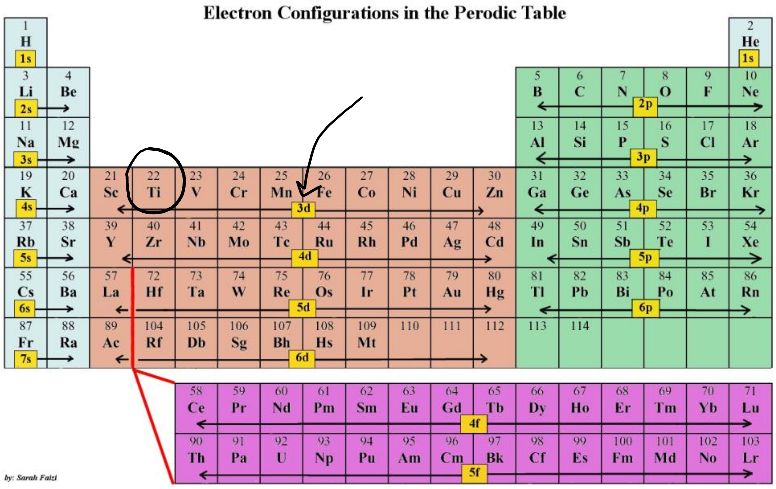
Energy Level Diagram.pptx

↿ clockwise spinning e-

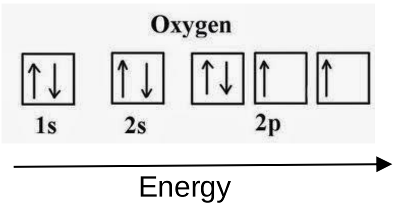
⇂ counterclockwise spinning e-

$_{22}Ti^0$

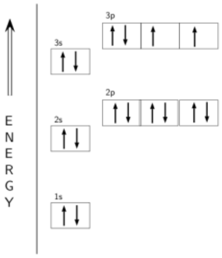




b) **Orbital (Block) Diagrams**



$_{16}S^0$



○ OR □ to represent the orbital

c) **** **Electron Configuration (complete/full and condensed/core/abbreviated etc…)**

How to Write Electron Configurations

PeriodicTable_e-config.png

$_{22}\mathrm{Ti}^{0}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

or

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$

complete

all superscripts must add to the atomic number for a neutral atom

Energy levels are arranged numerically (*textbook does it this way)*

[Ar] $4s^2 3d^2$

or

[Ar] $3d^2 4s^2$

condensed/abbreviated

Represents the same e- configuration as Argon, use the closest Noble gas that occurs ***before*** the element!

E.g.2

$_{50}\mathbf{Sn}^{0}$ $\mathbf{1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2}$ **[Kr]** $\mathbf{5s^2 4d^{10} 5p^2}$

Representation of Ions:

For **anions**, electrons are **added** in the order of **lowest** to **highest** energy level.
For **cations**, electrons are mostly **removed** from **highest *ns value***.

e.g.

ANIONS

$_{33}\mathbf{As}^{0}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$

$_{33}\mathbf{As}^{-3}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

3 more e-

Add 3e- since it gained 3e-
g

CATIONS

$_{23}\mathbf{V}^{0}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

$_{23}\mathbf{V}^{+2}$ $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3$

e- are removed from highest ns value

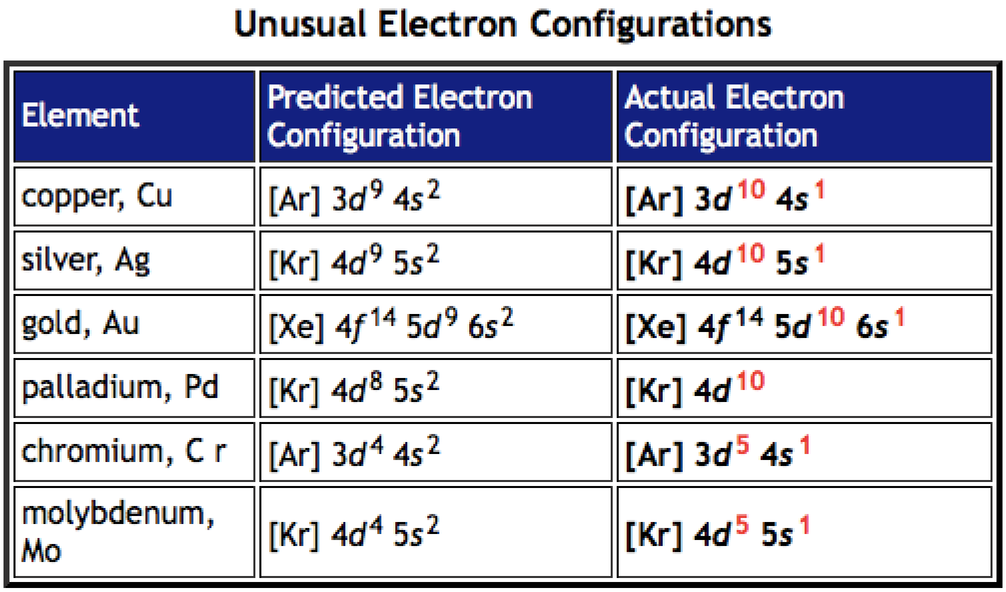
E.g. 2

$_{35}\mathbf{Br}^{-1}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$ $\mathbf{4p^6}$

$_{50}\mathbf{Sn}^{+4}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 4d^{10}$

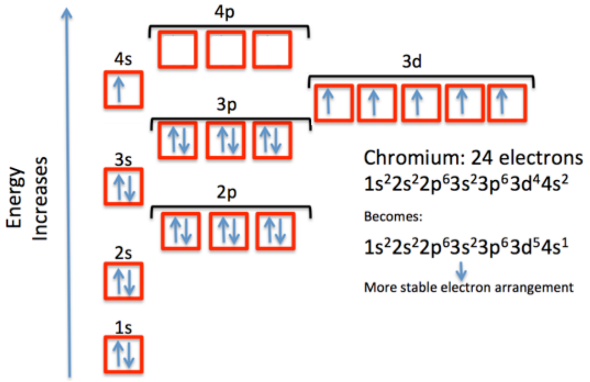
Exceptions to the Rules:

There are some exceptions to the e- configuration rules on the periodic table. We will focus on two of them in this course. **Half-filled or filled subshells** have an **extra stability** that may affect the electron filling in an atom, so if the subshell is **an electron away** from being **half filled** (**Cr, Mo**) or **filled (Cu, Ag, Au**) then **an electron** from the **s orbital** will be **promoted to the subshell**. Make sure you know these exceptions!



Unusual Electron Configurations

Element	Predicted Electron Configuration	Actual Electron Configuration
copper, Cu	[Ar] $3d^9\ 4s^2$	**[Ar] $3d^{10}\ 4s^1$**
silver, Ag	[Kr] $4d^9\ 5s^2$	**[Kr] $4d^{10}\ 5s^1$**
gold, Au	[Xe] $4f^{14}\ 5d^9\ 6s^2$	**[Xe] $4f^{14}\ 5d^{10}\ 6s^1$**
palladium, Pd	[Kr] $4d^8\ 5s^2$	**[Kr] $4d^{10}$**
chromium, C r	[Ar] $3d^4\ 4s^2$	**[Ar] $3d^5\ 4s^1$**
molybdenum, Mo	[Kr] $4d^4\ 5s^2$	**[Kr] $4d^5\ 5s^1$**



Aufbau Principle: Exception

NOTE: Elements that have the same number of electrons are called **ISOELECTRONIC.**

TOPIC 4: QUANTUM MECHANICS AND CHEMICAL BONDING

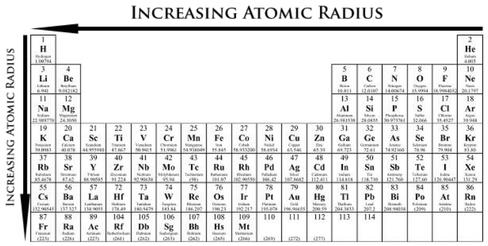
Review of Grade 11 Unit 1…

1. **Molecule** – an electrically neutral collection of atoms held together tightly enough to be considered as a single unit.

2. **Chemical bonds** – attraction between atoms within a molecule when atoms come together, electrons of each atom are under influence of nucleus & electrons of other atom

3. **Trends in the Periodic Table**

a) Atomic Radius



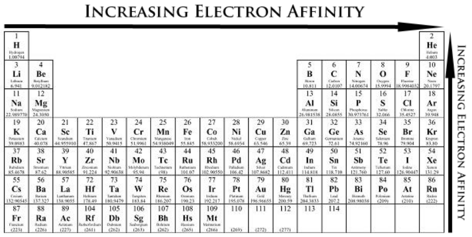
across a period, R→ L

-since are less valence e- there would be less EAF, therefore radius is larger

down a family

-since more n therefore more shielding and larger AR

b) Electron Affinity



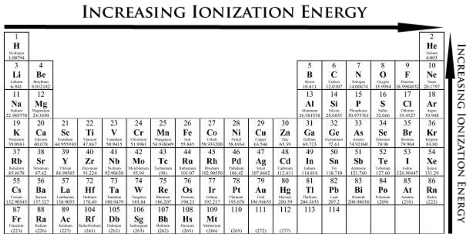
across L→ R

-since more valence e-; more EAF for outer shell, therefore free e- easily attracted

going up a family

- less n therefore more EAF for outer shell and free e- absorbed easily

c) Ionization Energy

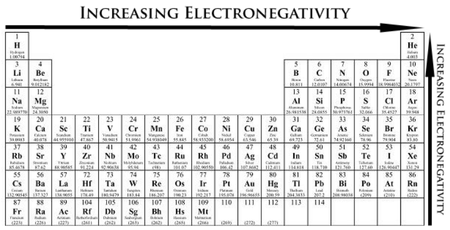


across L → R

-since more valence e- more EAF an harder to pull e- away. i.e. more IE needed

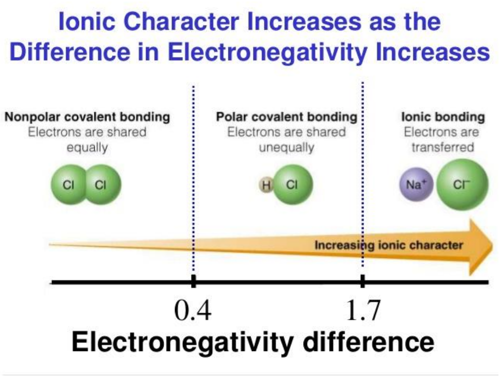
going up a family

-less n so more EAF so more IE needed to pull e- away



EN is the ratio of IE to EA its is numerical value that shows the Ability of an atom to hold onto a pair of e- when bonding.

e) Bonding Continuum ΔEN



4. Ionic Bonding-attraction between oppositely charged ions

e.g. **NaCl**

electron configuration: valence e^-

${}_{11}Na^0$ *low IE, EA)* $1s^2 2s^2 2p^6 3s^1$ ${}_{11}Na^{1+}$ $1s^2 2s^2 2p^6$

${}_{17}Cl^0$ *(high IE, EA)* $1s^2 2s^2 2p^6 3s^2 3p^5$ ${}_{17}Cl^{1-}$ $1s^2 2s^2 2p^6 3s^2 3p^6$

valence e^-

******NOTE: the valence electrons for the first 20 elements are the highest s & p e-**

Ionic equations:

Na^0 + IE -→ Na^{1+} + 1 e- Δ EN = 3.16-0.93= 2.23 (*ionic bond*)

Cl^0 + 1e- → Cl^{1-} + EA = $\boxed{NaCl}$

N.B. – Lewis dot diagrams are not typical drawn for ionic bonding

e.g. **Al_2O_3**

electron configurations: valence e^-

${}_{13}Al^0$ *(low IE, EA)* $1s^2 2s^2 2p^6 3s^2 3p^1$ ${}_{13}Al^{3+}$ $1s^2 2s^2 2p^6$

${}_{8}O^0$ *(high IE, EA)* $1s^2 2s^2 2p^4$ ${}_{8}O^{2-}$ $1s^2 2s^2 2p^6$

→ valence e^-

Ionic equations:

$(\ 2\ Al^0 + 2\ IE \rightarrow 2\ Al^{3+} + 6\ \cancel{3}e^-\ ) \times 2$ ΔEN = 3.44-1.61= 1.83 (*ionic bond*)

$(\ 3\ O^0 + 6\ \cancel{2}e^-\ \dashrightarrow 3\ O^{2-} + 3\ EA\ ) \times 3$ = $\boxed{Al_2O_3}$

Energy level (block) diagrams that show bonding

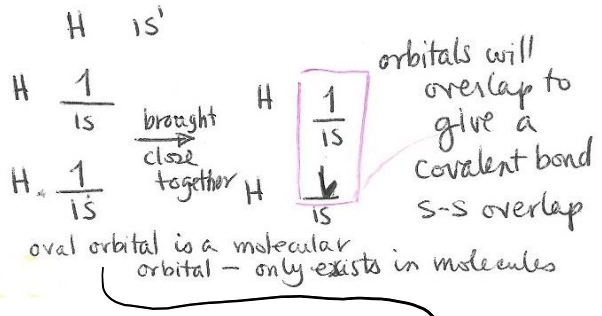
5. Covalent Bonding- sharing of e-

A) PURE COVALENT BOND-identical atoms sharing e-

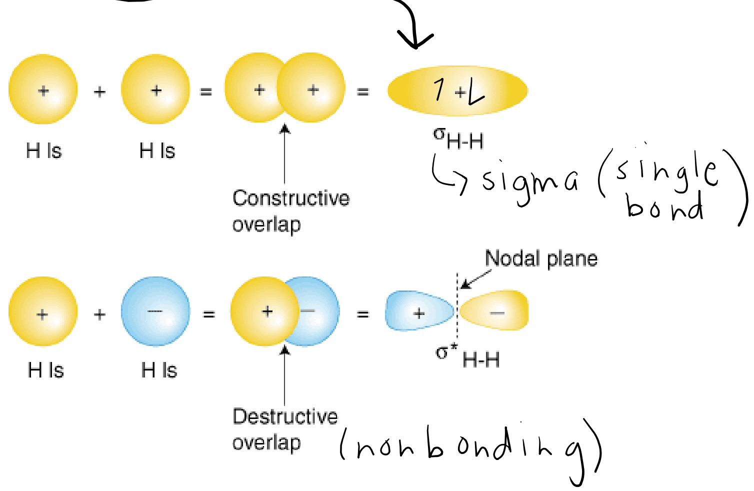
e.g. H_2

Lewis Dot Diagram:

Orbital (block) Representation:



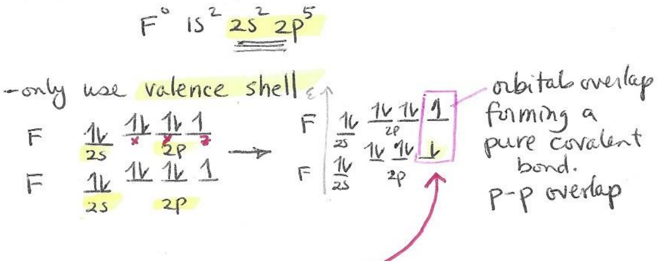
Molecular Orbitals:



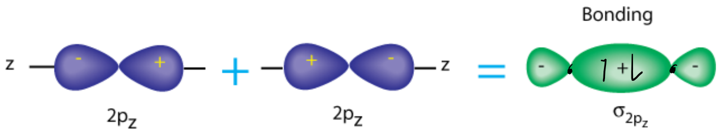
e.g. F_2

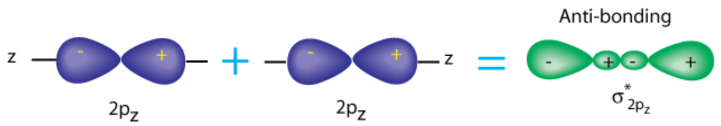
Lewis Dot Diagram:

Orbital Representation:



Molecular Orbitals:

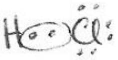




B) **POLAR COVALENT BONDS-unequal e- pair sharing**

e.g. HCl

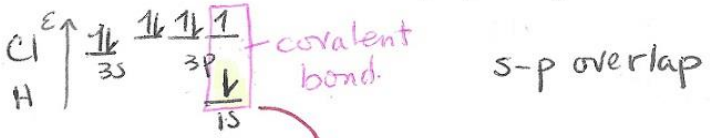
Lewis Dot Diagram:



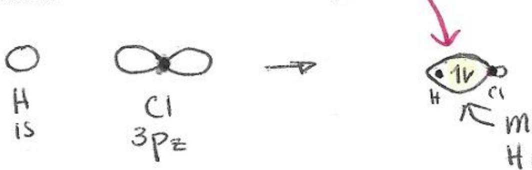
ΔEN 3.16
2.20
0.96

H 1s¹
Cl 1s² 2s² 2p⁶ 3s² 3p⁵

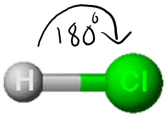
Orbital Representation:



Molecular Orbitals:

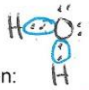


molecular orbital
H nucleus & Cl nucleus are at 180° from each other.



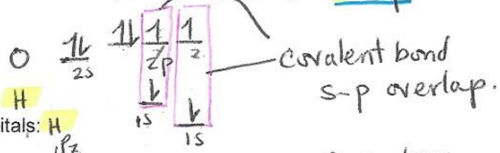
e.g. H_2O

Lewis Dot Diagram:

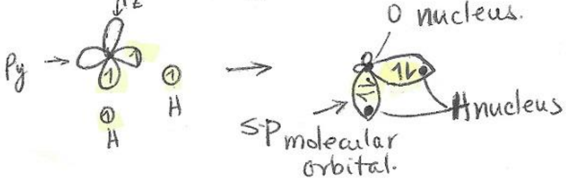


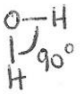
H 1s¹
O 1s² 2s² 2p⁴

Orbital Representation:

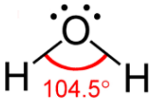


Molecular Orbitals: H





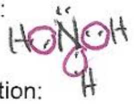
this structure is inconsistent with actual bond angle in H_2O



e.g. NH_3

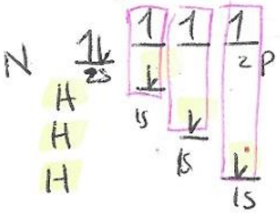
e.g. NH_3

Lewis Dot Diagram:

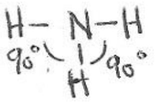


N 1s² 2s² 2p³
H 1s¹

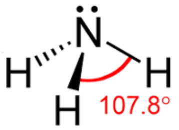
Orbital Representation:



Molecular Orbitals:



flat (2D)
but bond angle are incorrect



NOTE: for these last two, MOLECULAR ORBITALS CANNOT DESCRIBE THE ACTUAL 3-D STRUCTURE of the substance. As a result, HYBRID ORBITALS were developed. (TOPIC 6B)***

TOPIC 5A: LEWIS DOT/STRUCTURES DIAGRAMS

To draw a **Lewis dot diagram** the following rules are applied.

1. Count all **Valence electrons for each atom** and **add**. If the structure is a **cation** remove the same number of electrons as the charge, if the structure is an **anion** add the same number of electrons as the charge. Use these electrons in the manner below.

2. **Place least electronegative element in center**. This atom will be the central atom, the others will be the ligands (atoms attached to central atom).

3. Use **two electrons to bond each ligand to** the central atom.

4. Complete the **octet on each ligand except for hydrogen** (only give hydrogen 2 electrons). Hydrogen cannot go between 2 atoms, only around outside.

5. Place **any remaining electrons on the central atoms** as **lone pairs.**

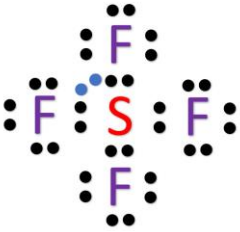
6. If the **central atom does not have an octet, form multiple bonds.**

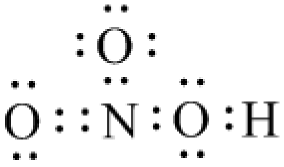
7. In **acids** with **oxygen, hydrogen is bonded to the oxygen.**
8. Calculate formal charges if applicable. (see Topic 6B)

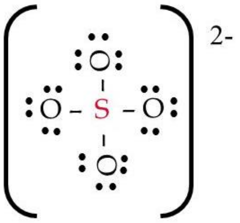
[Lewis dot diagrams.pptx]

e.g. **SF_4** **HNO_3** **SO_4^{-2}**

28e- **24e-** **32e-**







TOPIC 5B: WHEN LEWIS STRUCTURES FAIL; FORMAL CHARGES

- when following the rules in drawing **Lewis structures** for a molecule/ion and more than one possible Lewis structure exists, the method to determine which Lewis structure is **most feasible (more likely to exist in nature)** is to use **formal charges**

- **formal charge** is defined as the difference between the valence electrons in an isolated atom and the number of electrons assigned to the atom in the **Lewis structure**

- **in general,**

- a) Lewis structures with **no formal** charges are preferable to one showing formal charges;

- b) Lewis structures with **large formal charges** (+2, +3, -2, -3) are less plausible than those with smaller formal charges and

- c) in choosing between Lewis structures having formal charges, the most plausible is the one with the **negative charge placed on the most electronegative atom**

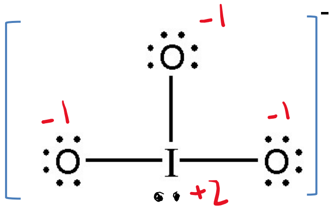
How to Calculate Formal Charges

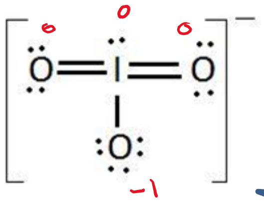
formal charge = total # of valence – non-bonding val e-1/2 bonding e-
OR
= total # of valance -lone electrons-bonds

- rules for use of formal charge:

1) for a neutral molecule, the **sum of formal charges must add up to be zero**
2) for **cations**, the **sum of formal charges must equal the charge on the ion**
3) for **anions**, the **sum of formal charges must equal the charge on the ion**

- **eg. IO_3^-**





More feasible ('likely') structure

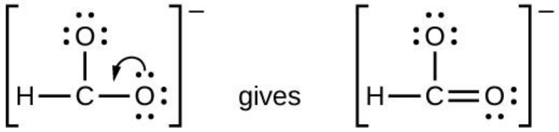
Lewis Structures and Formal Charges Practice

Note: You will only calculate formal charges in this course when specifically ask to do so when deciding between Lewis structures. Otherwise use the SIMPLE structure generated by using the rules from **Topic 5A** above.

TOPIC 5C: RESONANCE STRUCTURES

Resonance and Lewis Dot Structures

- there are some molecules/ ions where the Lewis structures drawn do not agree with experimental measurements of bond lengths and energies
- **eg.** CHO_2^-



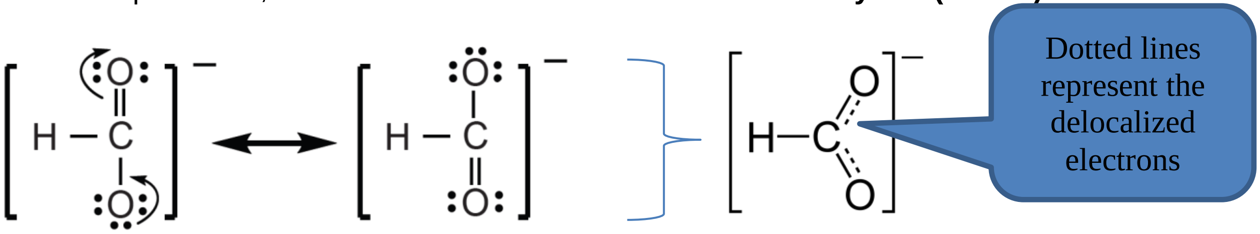
- from this Lewis structure, we expect that the data for bond lengths for the C-O bond would be: a C-O single bond and a C=O double bond length

- in fact, experiments show that the bond lengths are about ***halfway*** between the expected values for a single and double bond

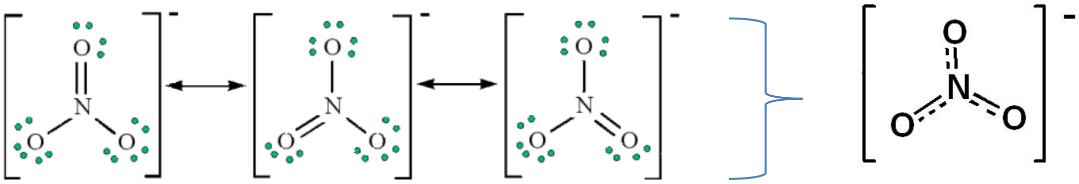
- to solve this problem, we use the concept of **resonance**

- we view the actual structure of the molecule or ion which we cannot draw satisfactorily as a composite or average of a number of Lewis structures

- for the example above, then: **Resonance Hybrid (Actual)**



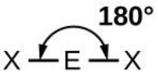
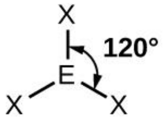
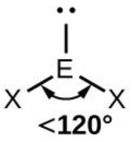
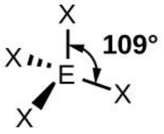
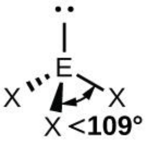
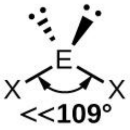
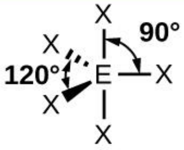
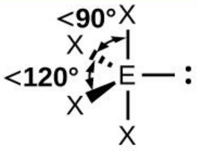
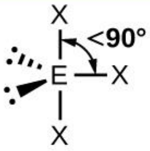
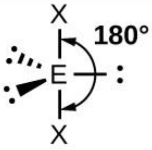
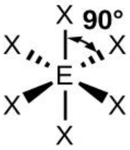
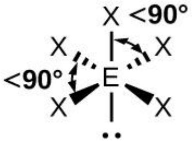
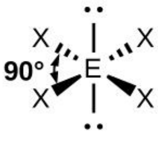
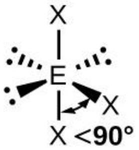
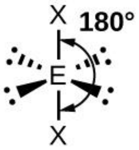
- these two Lewis structures are called **resonance structures** or **contributing structures** and the actual structure of the ion is said to be a **resonance hybrid** of two resonance structures
- the two-headed arrow is used to show that we are drawing resonance structures, not that the molecule moves between the two, but that the hybrid structure is a composite of the two
- to determine **how many resonance structures** a molecule/ion has**, examine the number of choices for the locations of double bonds, this number will give you the number of possible resonance structures**
- **eg.** NO_3^-



TOPIC 6A: Valence Shell Electron Pair Repulsion Theory

VSEPR is the **Valence Shell Electron Pair Repulsion Theory**. This theory was co-developed by Dr. Ronald Gillespie of McMaster University. VSEPR theory predicts the shape of simple molecules through the concept that valence electron pairs will repel each other, so, geometrically; these electron pairs will be as far apart as possible. In addition, the **degree of repulsion** varies:

Lone pair – lone pair repulsion > lone pair – bonding pair repulsion > bonding pair-bonding pair

Number of electron pairs	Electron pair geometries: 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 Linear				
3	 Trigonal planar	 Bent or angular			
4	 Tetrahedral	 Trigonal pyramid	 Bent or angular		
5	 Trigonal bipyramid	 Sawhorse or seesaw	 T-shape	 Linear	
6	 Octahedral	 Square pyramid	 Square planar	 T-shape	 Linear

EG: how the electrons are arranged around central atom

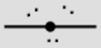
AX_nE_m Notation:

A= central atom
X= ligands (atoms attached to central)   (**n**)= # of ligands attached
E= electron pairs around central atom   (**m**) = # of lone pairs around central

Know this!

Molecular Geometry Summary Sheet

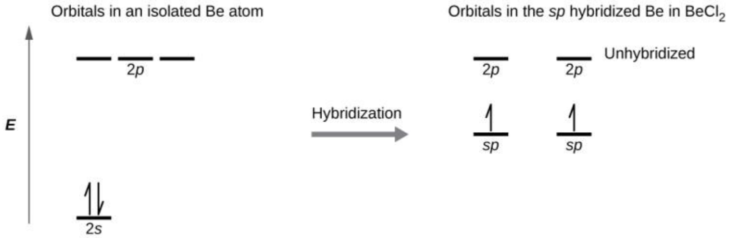
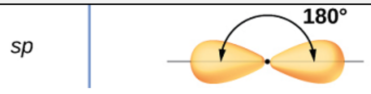
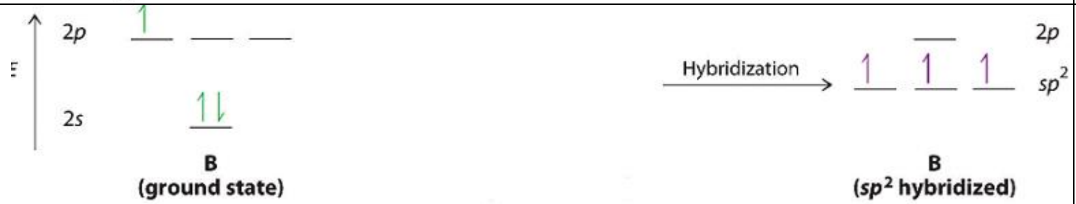
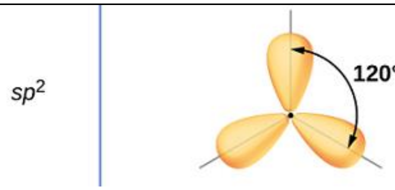
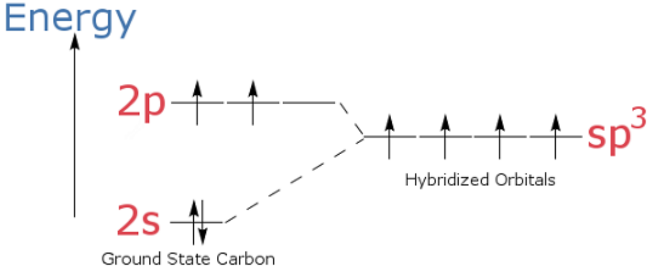
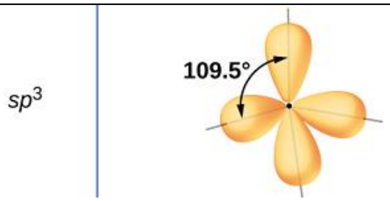
Shaded squares represent geometries which give non-polar molecules when all substituents X are identical.

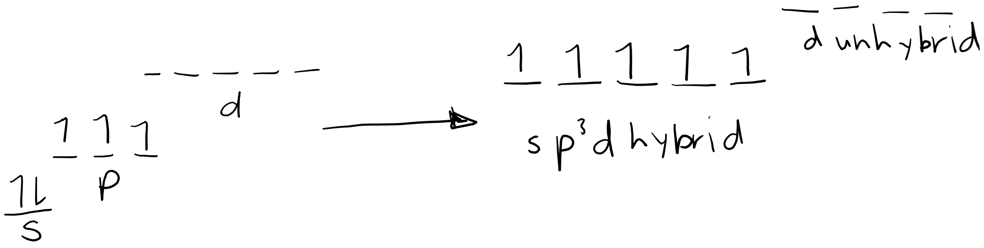
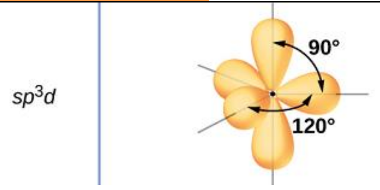
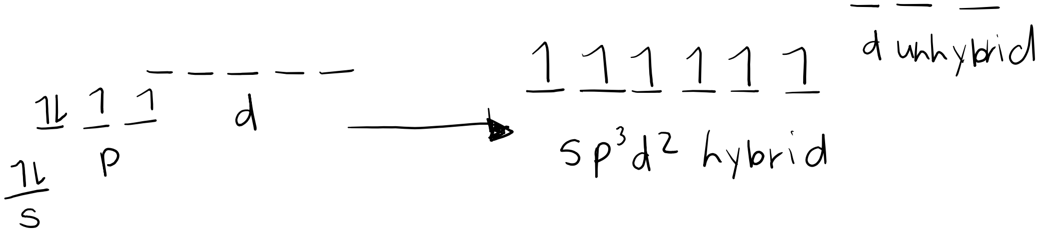
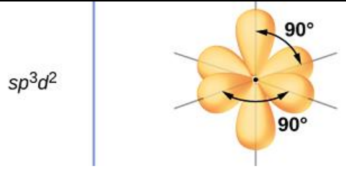
Total # of Groups of e^-	***Electron Pair Geometry*** *(Hybridization)*	***Approximate Bond Angle***	***# of Bonding Directions (# of X)***	***# of Lone Pairs (# of E)***	***Geometry Name*** *(VSEPR class)*	***Shape***	***Examples***
2	*linear* *(sp)*	*180^o*	2	0	**linear** (AX_2)		BeH_2, CO_2
3	*trigonal planar* *(sp^2)*	*120^o*	3	0	**trigonal planar** (AX_3)		BF_3, NO_3^-
			2	1	**bent** (AX_2E)		SO_2
4	*tetrahedral* *(sp^3)*	*109.5^o*	4	0	**tetrahedral** (AX_4)		CH_4
			3	1	**trigonal pyramidal** (AX_3E)		NH_3
			2	2	**bent** (AX_2E_2)		H_2O
5	*trigonal bipyramidal* *(sp^3d)*	*120^o (in plane) & 90^o (above & below)*	5	0	**trigonal bipyramidal** (AX_5)		PCl_5
			4	1	**seesaw** (AX_4E)		SF_4
			3	2	**T-shaped** (AX_3E_2)		ClF_3
			2	3	**linear** (AX_2E_3)		XeF_2
6	*octahedral* *(sp^3d^2)*	*90^o*	6	0	**octahedral** (AX_6)		SF_6
			5	1	**square pyramidal** (AX_5E)		BrF_5
			4	2	**square planar** (AX_4E_2)		XeF_4

TOPIC 6B: HYBRID ORBITALS AND VSEPR THEORY

Covalent bonding can also be explained using **hybrid orbitals**. These are the **mixing of two or more of the basic atomic orbitals** of an atom when bonding occurs forming molecular orbitals). These **new orbitals** have **new shapes** and directional properties thus forming **stronger more stable** (minimum E_p) than would be possible if just simple atomic orbitals were used.

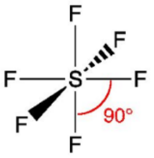
Hybridization

NAME OF NEW HYBRID ORBITALS (*steric #, & # of new orbitals formed)	**ATOMIC ORBITALS mixed TOGETHER** (valence shell of central atom) → Forming **HYBRID ORBITALS**	**NEW HYBRID ORBITAL SHAPE (aligns with VSEPR Theory)**	**E.g**
sp * 2			$BeCl_2$, BeH_2
sp² *3			BH_3, AlH_3 BCl_3 $SnCl_2$
sp³ *4			CH_4 NH_3 H_2O

sp^{3}d *5		 ***Any element that is in n≥3 can form more then an octet, has d orbitals	PCl$_5$, SF$_4$ ClF$_3$ XeF$_2$
sp^{3}d^2 *6			SF$_6$ IF$_5$ XeF$_4$

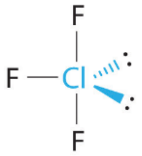
Note: ***Steric Number(domains) = number of bonding atoms plus lone pairs around the central atom in a molecule. Double and triple bonds are counted as 1 electron pair/group. Steric number indicates the hybridization, and how many hybrid orbitals of that type are formed when bonding. Steric numbers are also used to predict VSEPR shapes.

For Example: SF$_6$, ClF$_3$



S 6 atoms bonded + no lone pairs = (steric # of 6) = 6 sp^{3}d^2 hybrid orbitals formed to bond

Cl 3 atoms bonded + 2 lone pairs = (steric # of 5) = 5 sp^{3}d hybrid orbitals formed to bond



TOPIC 6C: SIGMA AND PI (π) BONDING (*Different types of orbital overlap in covalent bonding*)

Sigma and Pi bonding

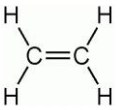
Sigma bonding (σ) – in single bonding **(s, p and hybrid orbitals) head-on overlap**

- electron density is concentrated between the two bonding nuclei (i.e. orbitals overlap between the two bonding nuclei)
- accounts for the framework and geometry of the molecule
- **enables rotation without bond breaking**

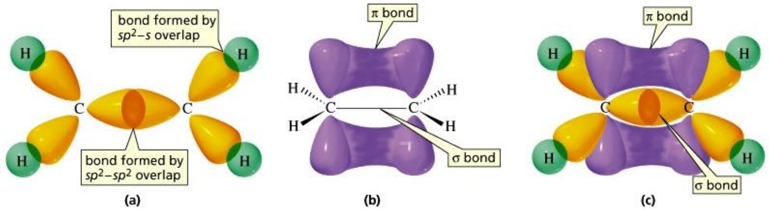
Pi bonding (π) – found in double and triple bonding **(unhybridized orbitals) side to side overlap**

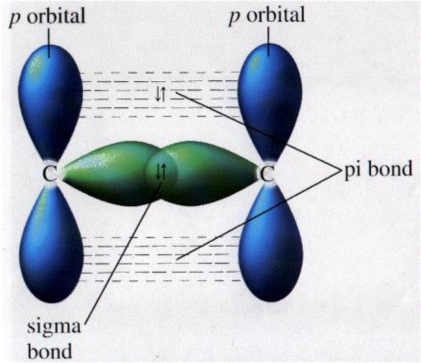
- electron density is concentrated in **two** separate regions that lie on opposite sides of an imaginary line joining the two nuclei
- provides for multiple bonding (double and triple bonding) where the bond occurs in a different plane to the sigma bond
- note – a double bond consists of a sigma and pi bond while a triple bond consists of a sigma and two pi bonds

e.g.

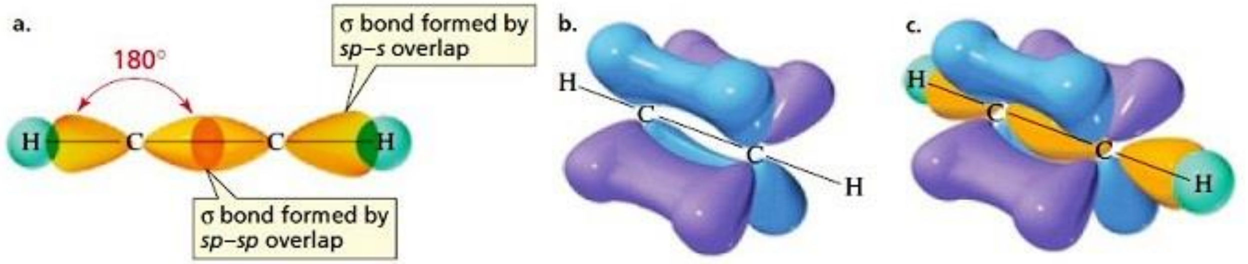


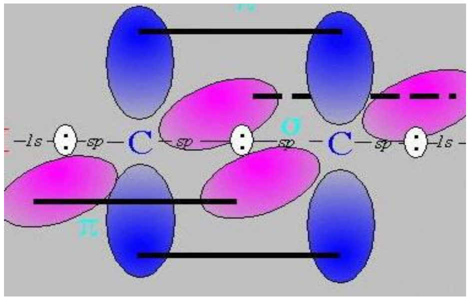
Ethylene (Ethene)





Acetylene (Ethyne), $H-C\equiv C-H$



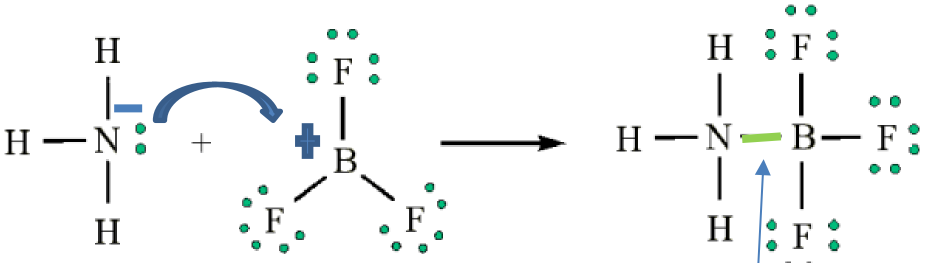


pi bonding can be delocalized, that is the electrons involved in the pi bonding can spread among the nuclei which are bonded

TOPIC 7: CO-ORDINATE COVALENT BONDING

Recall: when **both electrons** of the shared pair come from **one** of the two atoms

e.g. #1 NH_3 + $BF_3 \rightarrow$ NH_3BF_3

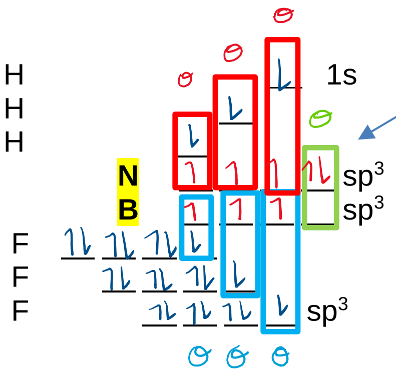


N is sp^3

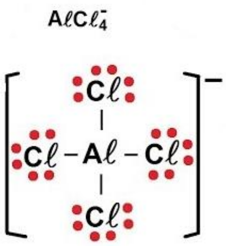
B is sp^2, there is an empty p orbital

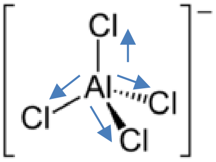
both N and B are sp^3

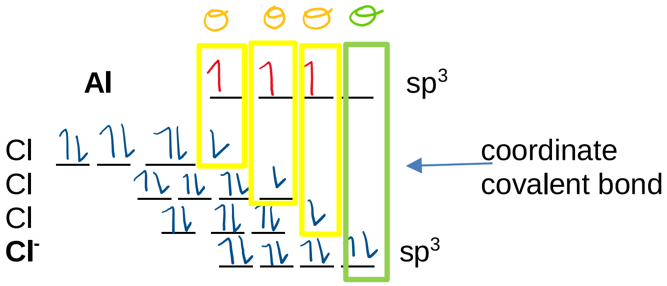
coordinate covalent bond



e.g. #2 – For the following substance **$AlCl_4^-$**, draw:
a) the Lewis Dot Diagram and indicate the orbital hybridization on Al
b) AXE notation, 3-D VSEPR shape and name both the electronic and molecular geometries
c) Molecular polarity of the molecule
c) the bonding block diagram, identify the coordinate covalent bond







Al: sp^3

AX4
EG: tetrahedral
MG: tetrahedral
M polarity: n/a, an ion

TOPIC 8A: AGGREGATES: BONDING AND PROPERTIES

Aggregates are groups of atoms, molecules or ions that are held together.

Two types of aggregates exist: _**amorphous**__and **crystalline**_____________

Amorphous solids, like glass, plastic and polymers are **covalently bonded network** with ______**little**_________ ordering.

Crystalline solids have:

- specific structural units that make up the solid and
- bond or a force that exits between the units to hold them in place within the crystal structure.

Two types of forces exist:

INTRAMOLECULAR FORCE

- electrostatic force of attraction **WITHIN** a molecule
- has stored energy between 200 – 800 kJ/mol.

INTERMOLECULAR FORCE

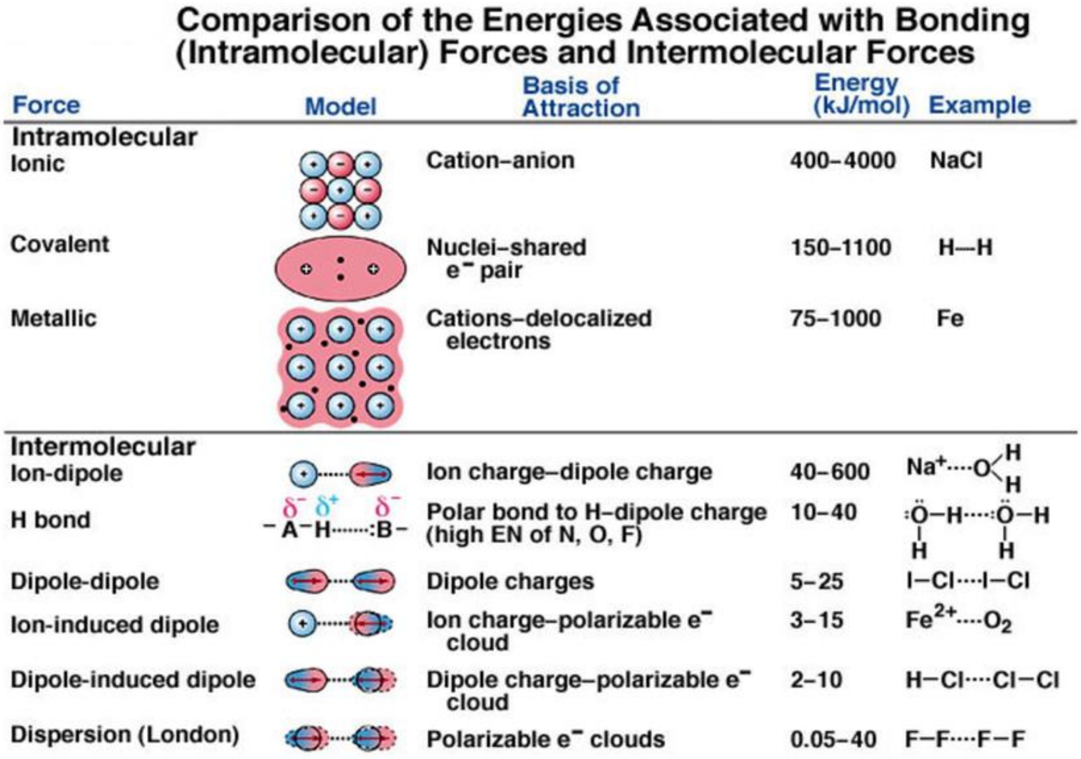
- electrostatic force of attraction **BETWEEN** particles
- has stored energy of 125 kJ/mol

The physical properties of substances, such as melting point, boiling point, conductivity, luster and viscosity, depend on:

- the structural units making up the aggregate (kind of particles the substances are made up of)
- bond or force that exists **BETWEEN** the units (INTERMOLECULAR FORCES)

Intermolecular forces found in crystals can be divided into STRONG and WEAK

Strong forces are:	Weak intermolecular forces are:
► **metallic bonding**	► **London dispersion force**
► **ionic bonding**	Vander Waals forces
► **covalent bonding**	► **dipole-dipole forces**
	► **hydrogen bonding**



Comparison of the Energies Associated with Bonding (Intramolecular) Forces and Intermolecular Forces

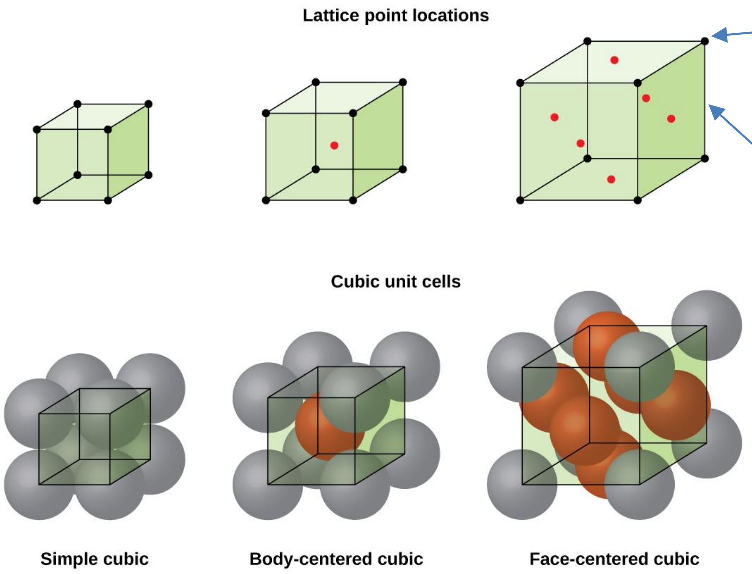
Force	Model	Basis of Attraction	Energy (kJ/mol)	Example
Intramolecular				
Ionic		Cation–anion	400–4000	NaCl
Covalent		Nuclei–shared e^- pair	150–1100	H—H
Metallic		Cations–delocalized electrons	75–1000	Fe
Intermolecular				
Ion-dipole		Ion charge–dipole charge	40–600	$Na^+\cdots O\genfrac{}{}{0pt}{}{H}{H}$
H bond	$\overset{\delta^-}{-A}-\overset{\delta^+}{H}\cdots\cdots:\overset{\delta^-}{B}-$	Polar bond to H–dipole charge (high EN of N, O, F)	10–40	$:\underset{H}{\underset{\mid}{\ddot{O}}}-H\cdots\underset{H}{\underset{\mid}{\ddot{O}}}-H$
Dipole-dipole		Dipole charges	5–25	I—Cl····I—Cl
Ion-induced dipole		Ion charge–polarizable e^- cloud	3–15	$Fe^{2+}\cdots O_2$
Dipole-induced dipole		Dipole charge–polarizable e^- cloud	2–10	H—Cl····Cl—Cl
Dispersion (London)		Polarizable e^- clouds	0.05–40	F—F····F—F

There are four different types of crystalline solids:

TYPE OF SUBSTANCE (crystalline solid)

IONIC	**MOLECULAR**		**NETWORK (Covalent)**	**METALLIC**
	Polar	Non-polar		
Ionic Bonding	LD & Dipole Forces Hydrogen Bonding	LD forces	Covalent Bonding *(can have LD)*	Metallic Bonding
NaCl	H_2O	Cl_2, CH_4	diamond	aluminum

Crystalline solids consist of patterns of repeating unit cells called the crystal lattice. Crystal lattice structures are studied extensively. The structure enables the transfer of energy.



Lattice points-can be an
-atomic kernel
-non-polar molecule
-polar molecule
-atom
-ion

Represents the **attractive force** that's hold the lattice points together, i.e.
-metallic bonding
-IMF's LD
-IMF's LD, DD, HB
-LD
-ionic bonding

TOPIC 8B: BONDING WITHIN CRYSTALS

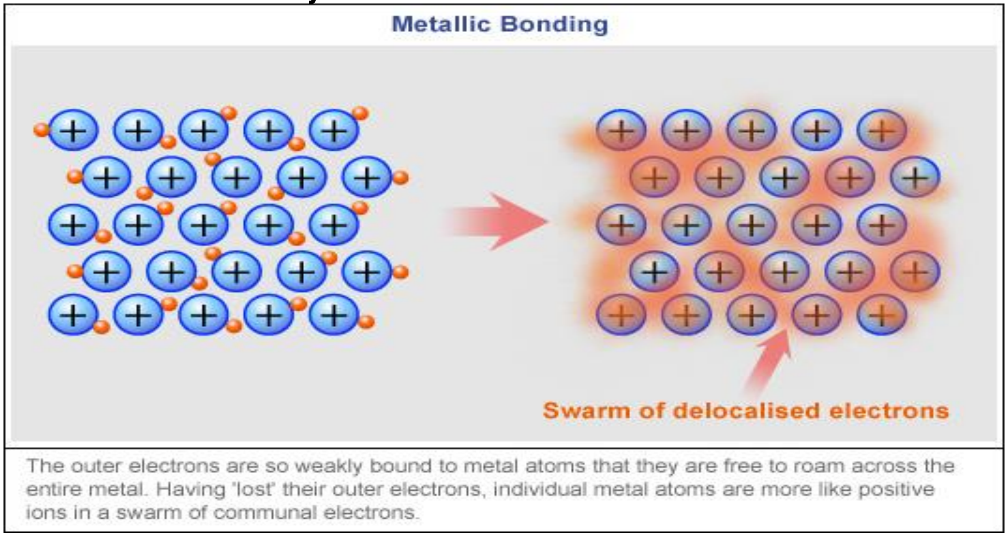
The Strong Forces are:

a) Ionic Bonding

- bonding is the electrostatic attractive force between oppositely charged ions

b) Metallic Bonding

- is the electrostatic attraction that exists between the **valence electrons of one atom** and the **atomic kernel** of an **adjacent atom**



The outer electrons are so weakly bound to metal atoms that they are free to roam across the entire metal. Having 'lost' their outer electrons, individual metal atoms are more like positive ions in a swarm of communal electrons.

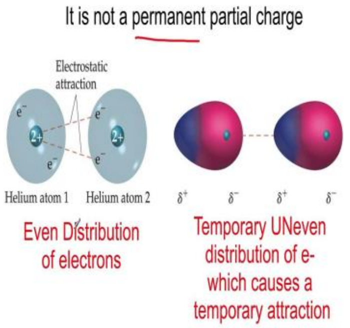
c) Covalent Bonding

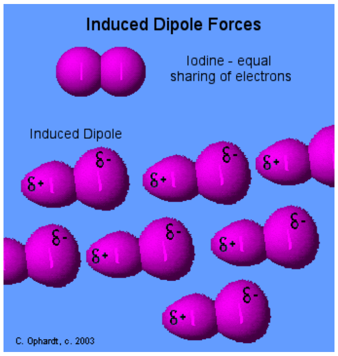
- is the simultaneous attraction of a pair of electrons for two different nuclei

The Intermolecular Forces are:

a) London Dispersion Force

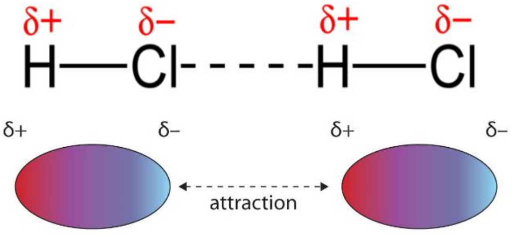
- result from instantaneous non-permanent dipoles created by random electron motion.
- are present in all molecules and are **directly proportional to molecular size** - **i.e. total number of electrons and surface area**





b) Dipole-Dipole Forces

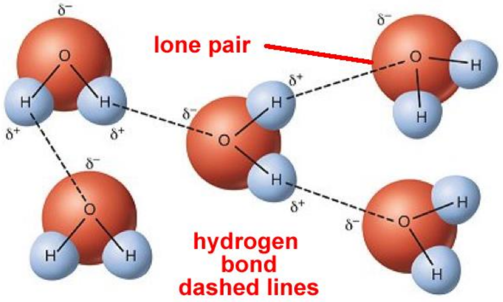
- result from the electrostatic attractive force that exists in **polar molecules** (i.e. molecules with permanent dipoles)



ΔEN = 3.16-2.20 = 0.96
MP: Polar Molecule

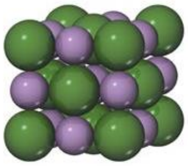
c) Hydrogen Bonding

- attractive forces between two molecules that have within the molecule **hydrogen bonded** to **N, O or F** (these atoms have the highest electronegativity)
- the bond is very polar and each of these atoms has lone pairs to which the hydrogen atom of an adjacent molecule will become attracted



Topic 9A: Ionic Crystals

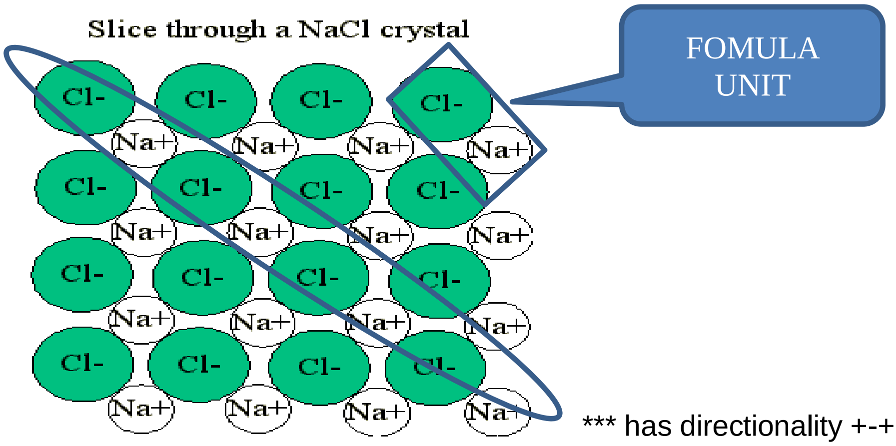
Ionic Solids



Ionic solids

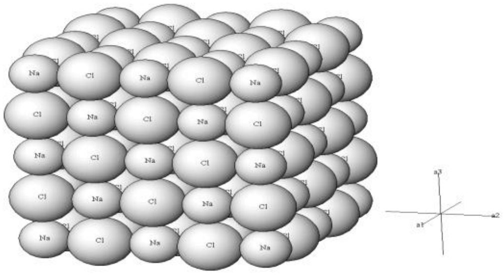
Extended networks of ions held together by ion–ion interactions (NaCl, MgO)

- made up of positive and negative ions held together by electrostatic forces between oppositely charged ions. These **bonds** are **strong** and **directional**
- example – NaCl (table salt) is made up of Na^{1+} and Cl^{1-}
- In **2 – Dimensions:** alternating pattern of Na^{1+} and Cl^{1-} ions, held together by strong electrostatic forces



*** has directionality +-+-

- In **3-Dimensions**: each type of ion forms a face centered cubic arrangement



***Lattice point is an ion.

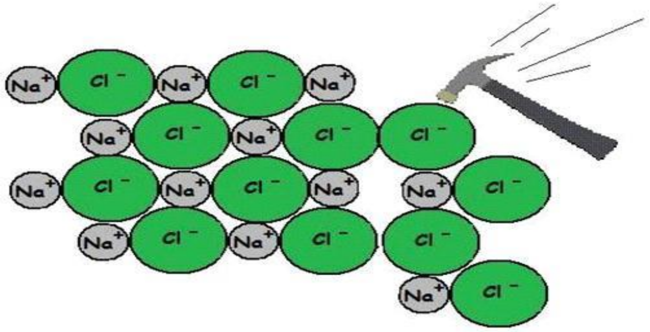
Properties of Ionic Substances:

- **Hard Solids –** the attractive forces between oppositely charged ions are **very strong**
- **High melting point –** need great amount of energy to break the strong electrostatic attractive forces between oppositely charged ions**; the smaller ions form stronger forces** (ions are closer together) therefore more stable hence **higher melting point and lower solubility**

Substance	Melting point (° C)
NaF	988
NaCl	800
NaBr	790
NaI	660

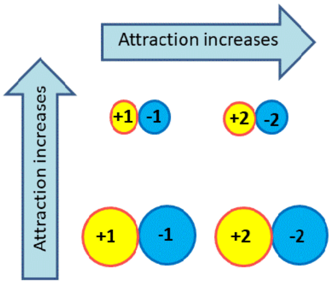
The F^- has a smaller radius so can pack more closely & have stronger ionic bonding

- **Brittle –** force of attraction between ions is directional; if force is applied, **ions shift in position**, causing **like charges** to be near each other → **REPULSION** and crystal shatters



Repulsion between close like charges resulting in the crystal shattering i.e. The directionality is destroyed.

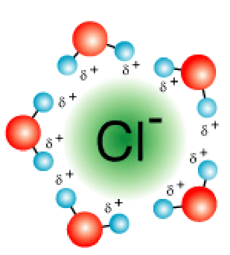
- **Electrical conductivity**
  does not conduct in the solid state; excellent conductors in **molten state** or when **dissolved in water** this is because electrostatic force of attraction between ions is very strong in the solid state and the charged particles **are not free to move**, whereas in the molten state or when dissolved, the force between ions is overcome and the ions are free to move



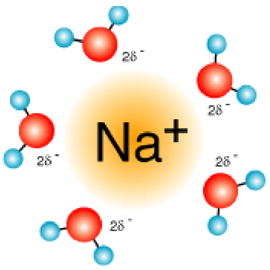
- **Non-volatile (does not evaporate into a gas easily= no/low odour)**
  Electrostatic attractive force between ions is strong – note that the **greater the charge** of the ions the **stronger the force**

- **Soluble in water**
  Since **water is polar**, the **dipoles** will react with the **poles of the ionic substance** resulting in dissociation and the ions becoming aquated.

  Eg. $NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$



Slightly positive hydrogen are attracted to chlorine anions

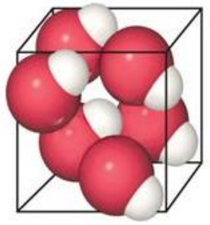


Slightly negative oxygen are attracted to sodium cations

TOPIC 9B: MOLECULAR CYRSTALS

Molecular Solids

- Covalent molecules occupy the lattice points of the crystal lattice to form solids
- Examples **H_2O (ice)**, **CO_2 (dry ice), I_2 (solid iodine)**
- The properties of these solids **depend on polarity, shape** and arrangement of the basic structural units. These crystals are divided in **polar** and **non-polar** molecular solids



Molecular solids

Discrete molecules held together by intermolecular forces (HBr, H_2O)

Properties of Molecular Crystals

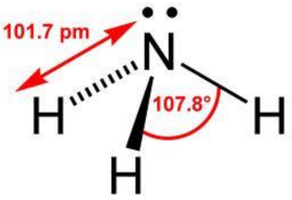
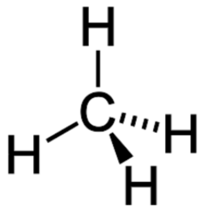
1. Liquids and solids are **NOT** very good conductors.
   Reason: No charged particles: lattice points are occupied by neutral particles (molecules)

2. Many exist as **gases at room temperature**; they have a **low melting point** and usually **low boiling points**. **Solids are soft and may have a waxy consistency**
Reason: **intermolecular forces are weaker than inter-ionic forces** thus forces between neutral atoms or molecules are fairly weak

3. Needs lots of energy to decompose these solids
   Reason: intramolecular forces > intermolecular forces
   E.g. $H_2O\ (s) \rightarrow H_2O\ (l)$ at 0 °C (*breaking IMF*)
   $H_2O\ (l) \rightarrow H_2O\ (g)$ at 100 °C (*breaking IMF*)
   $H_2O\ (l) \rightarrow H_2\ (l) + \frac{1}{2} O_2\ (g)$ at > 2000 °C (*breaking covalent bonds*)

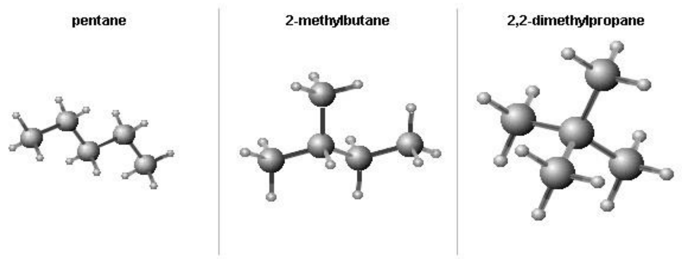
4. Variations exist between molecular crystals due to **polarity and molecular shape**

   a) Polarity - intermolecular forces between **polar molecules** are **stronger** than between **non-polar molecules**

NH_3 mass =17 u; m.p.= - 77 °C	**CH_4 mass =16 u; m.p.= - 182 °C**
 MP: polar (10e-) LD DD HB More forces the higher MP	 MP: non-polar (10 e-) LD only

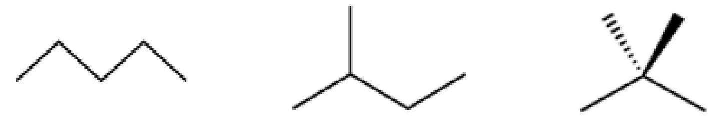
b) Molecular Shape affects the m.p. and b.p. of these crystals

e.g. consider pentane – **C_5H_{12}** which has three possible structural isomers: (trivial names: n-pentane, iso-pentane and neo-pentane)



All have the same # of electrons
-all are non-polar molecules, but they don't have the same mp/bp!

Skeletal diagram



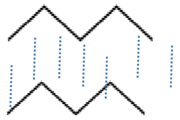
MP (°C)	−129.8	−159.9	−16.6
BP (°C)	36.0	27.7	9.5

Straight Chain

Branched chains

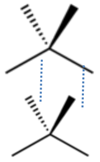
<u>In liquid form: (boiling)</u>

In the straight chain, molecules are closer together with a **large surface area**. This causes the forces of attraction between molecules to **increase with a resulting boiling point increase**. In the **branched chains**, **less surface area exists** resulting in less attractive force between molecules and a **decrease in boiling point.**
-no crystals and therefore no lattice point in the liquid forms



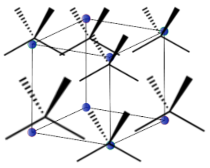
- lots of surface area
-lots of LDFs

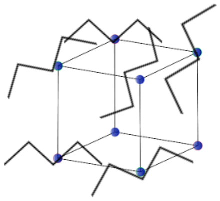
Small surface area, less LDFs



<u>In solid form: (melting)</u>

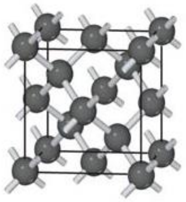
In the **straight chain**, the individual molecules **do not fit easily into lattice points** causing an **unstable crystal arrangement** hence a **lower melting point since it is easy to separate** the molecules. In **the branched chain** of neo-pentane, the molecule is able to **fit easily into the lattice** and forms a **stable crystal**. This isomer has a **higher melting** point and the **molecules are difficult to separate.**





TOPIC 9C: COVALENT / NETWORK CRYSTALS

Network Solids



Covalent-network solids

Extended networks of atoms held together by covalent bonds (C, Si)

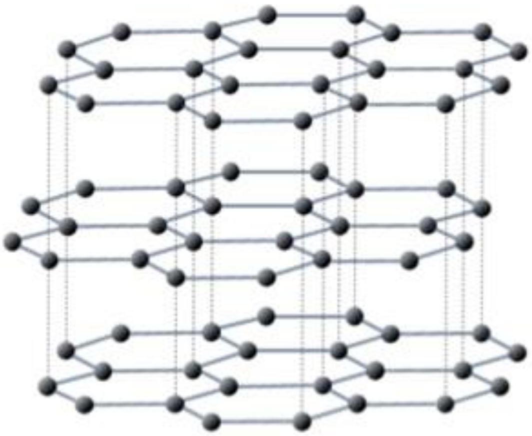
- a 1-D, 2-D or 3-D arrangement of covalent bonds that extends throughout the crystal
- this yields a gigantic continuous **network** of atoms held together by **single covalent bonds**
- can be of the **same element** or **different elements** (i.e. crystal lattice can be the same atom or a different atom)
- this forms substances with **higher melting points and greater hardness than salts (ionic)**

e.g. quartz (SiO_2) asbestos (SiO_4), carbon: diamond, graphite, or fullerenes (e.g. buckyball)

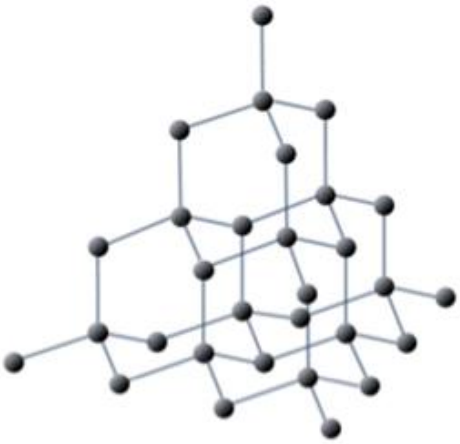
Diamond, graphite and fullerenes are allotropes of carbon – i.e. different dimensional arrangement of the carbon atom.

Diamond – 3-D structure of sp^3 hybridized carbon atoms; this results in a **very hard solid** since each carbon atom is covalently bonded to four other carbon atoms in a tetrahedral structure; it is non-conductive since there are no delocalized electrons and has an **extremely high melting point** (**above 3500 °C**); it is very dense since the carbon atoms can pack very closely together; the **bonding is very directional resulting in diamonds cleaving (shattering) under stress.**

Graphite – **2-D structure of sp^2 hybridized carbon atoms**; this results in layers of carbon atoms; within the layers are delocalized electrons **resulting in the conductivity ability of graphite**; holding the layers together is **London Dispersion force** which accounts for the **softness of graphite** (i.e. the layers slide off of each other); it has a **high melting point due to the bonding between the carbon atoms in each layer.**

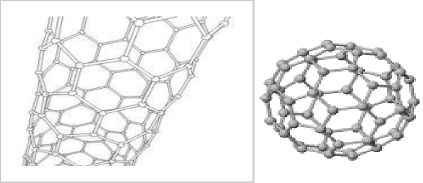


Graphite (solid lines are strong covalent bonds, dotted lines are weak inter-layer bonds)



Diamond (all bonds are strong covalent bonds)

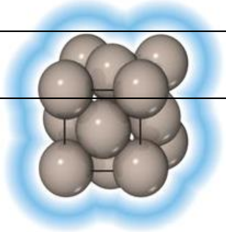
Fullerenes - buckyball (discovered in 1986) have since discovered number of variations -- e.g. tubules etc.



One Dimensional Network Solids:

- consists of long chain of molecules bonded covalently
- each chain held together by **weak intermolecular forces**, giving this type of solid substance a “fibrous” structure
- most **polymers f**all into this category

TOPIC 9D: METALLIC SUBSTANCES



Metallic solids

Extended networks of atoms held together by metallic bonding (Cu, Fe)

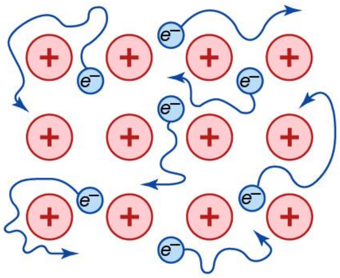
Metallic Solids

TYPE OF ELEMENT	BOND TYPE	ELEMENT	BONDING ORBITALS AFTER PROMOTION
Metal	Metallic	Elements with few valence electrons	s ___ sp ___ ___ sp^2___ ____ ___
Non-Metal	Network	Elements in family 14	sp^3 ___ ____ ___ ____
	Molecular	All other non-metals	Families 15, 16, 17

Metallic elements have: **few valence** electrons, **vacant valence orbitals** and **low ionization energies.**

Properties of Metallic Solids

metallic bond



Metallic crystals consist of a **3-D**, closely packed network of **positive ions** surrounded by a sea of **delocalized mobile electrons**. The electrons can move uninhibited throughout the crystal. The **electrons are held** within the metal by the attraction of the **positive kernels** and the positive kernels are held together by the electrostatic attraction of the electrons that move between them.

As a result, **metals:**

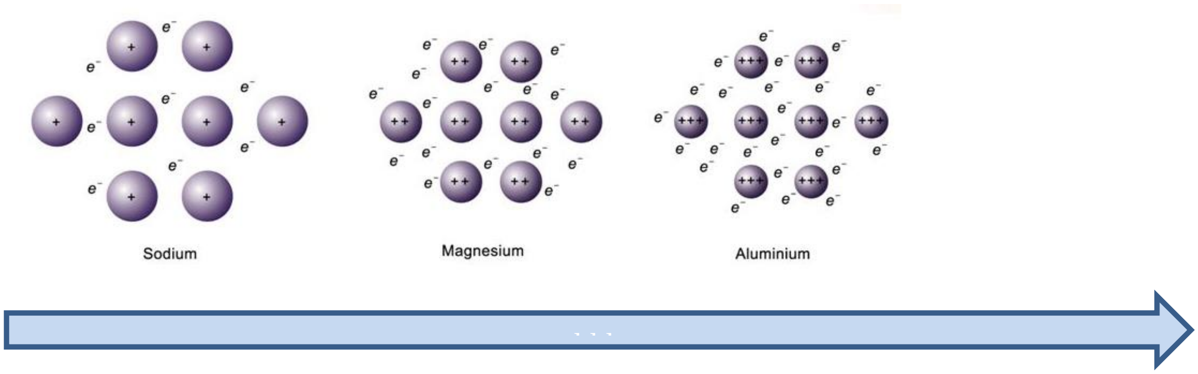
- are bright, shiny and silvery-white (exceptions: Au, Cu and Bi)
- are very conductive: able to conduct both electricity and heat
- are malleable and ductile
- are soft or hard
- have high melting and boiling points (wide variation)

Reasons for these properties:

- **luster:** light energy causes the **valence electron to oscillate** and radiate energy at the same frequency as the incident light
- **electrical conductivity: mobile electrons move easily between the atomic kernels** when an applied current is used
- **heat conductivity: electrons move from regions of high temperature where** they attain high kinetic energy and transfer it to the **cooler regions of the crystal lattice**. Heat, is transported by vibrational motions of the particles making up the crystal.
- **Malleability and ductility: metals can easily be worked because they do not have fixed bonds. The metallic bonds are non-directional** which means that the metal crystal lattice can easily be deformed **without fracturing or shattering**.

The* strength of the metallic bond depends on three main factors:

- **nuclear charge** (number of protons)
-**number of valence electrons**
-**size (atomic radius)** of the metal atom



Increasing strength of metallic bonding

As the number of **valence electrons increases**, the **strength of the metallic bond increases.** In general, as the strength of the metallic bonds increases, **the melting point, boiling points and hardness increase**. **Transition metals** should have **the highest melting and boiling point and** be the **hardest** because they possess **partially filled d and s orbitals** with **more** delocalized electrons.

Crystal Type	Particles in Crystal (Lattice pts)	Principal Attractive Forces between Particles	Melting Point	Electrical Conductivity of Liquid	Characteristics of Crystal	Conditions for formation	Examples
IONIC CRYSTALS	Positive and negative ions	Electrostatic attractions between ions. (Ionic Bonding) Very strong 600 – 4000 kJ/mol	High	High But can't conduct as a solid	Hard, brittle. Most dissolve in polar solvents (when dissolved can conduct)	Formed between atoms of widely differing electronegativity	KCl, CaF_2, Cs_2S, MgO
COVALENT NETWORK CRYSTALS	Atoms	Covalent bonds. Very strong 300 – 800 kJ/mol -may have LD	Very High	Depends on type. 2- D will conduct in solid state because of the delocalized e-	Very hard & dense Insoluble in most ordinary liquids	Most are formed by two elements of group 14 or by elements whose average periodic group number is 4	Diamond, graphite, SiC, AlN, BeO, $CuCl_2$ SiO_2 (quartz) SiO_4 (asbestos)
METALLIC CRYSTALS	Positive metal atom kernels plus mobile electrons	Metallic bonds. Strong 50 – 800 kJ/mol	Most are High (depends on number of valence e-)	Very High delocalized valence e;	Most are hard, malleable, ductile. High thermal(heat) conductivity. Generally insoluble in liquids. Usually soluble in molten metals.	Formed by electro positive elements (low electronegativity) Few valence e-	Cu, Ca, Bi, Pb, Zn, La, Fe, V, Li, Cr, Pt, CuZn(alloy)
MOLECULAR CRYSTALS POLAR	Polar molecules	Electrostatic attraction between dipoles. Intermediate strength. Can be strengthened by LD, DD and hydrogen bonding	Intermediate	Very low	More fragile than ionic crystals. Most are soluble in polar solvents. -They can ionize to form ions & conduct	Formed from unsymmetrical molecules containing polar covalent bonds. Such bonds are formed between atoms having a moderate difference in electronegativity	PF_3, $CHCl_3$, ICl, CH_3COOH, H_2O, CO, NH_3, SO_3, many organic compounds
MOLECULAR CRYSTALS NON-POLAR	Atoms or non-polar molecules	Van der Waals forces Weak. 0 – 50 kJ/mol (LD)	Low	Extremely low	Very soft. Most are soluble in non-polar or slightly polar solvents	Formed from atoms or from symmetrical molecules containing only low polarity covalent bonds. Such bonds are formed between like atoms or atoms having a small difference in electronegativity	Ar, H_2, Cl_2, S_8, C_6H_6, CH_4, CO_2, I_2, CCl_4, P_4O_{10}, Ne, BF_3

CHARACTERISTICS OF CRYSTALINE SOLID