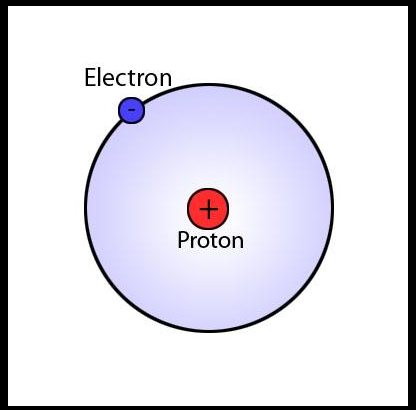
Unit 11 Outcomes

LPT

Describe the extra-nuclear structure of the atom



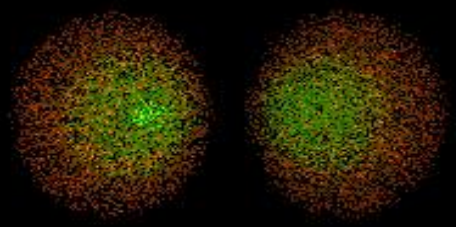
Bohr Model

Electron is a solid particle that revolves around an orbit
Orbit is defined by laws of magnetic attraction
Only works for single electron (hydrogen)

Quantum Model

Electron behaves as a particle-wave (a duality)
Electron position & path cannot simultaneously be determined
Electron described by $H \Psi = E \Psi$ where
Ψ^2 gives probability of finding an electron in space
3-D plot of Ψ^2 gives regions of electron occupancy

Ψ^2 Plot of P Orbital



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Give the qualitative relationship between:

wavelength and frequency

Frequency & Wavelength are inversely related:

high frequency means short wavelength

low frequency means long wavelength

wavelength and energy

Wave energy & wavelength are inversely related

wavelength increases, energy decreases

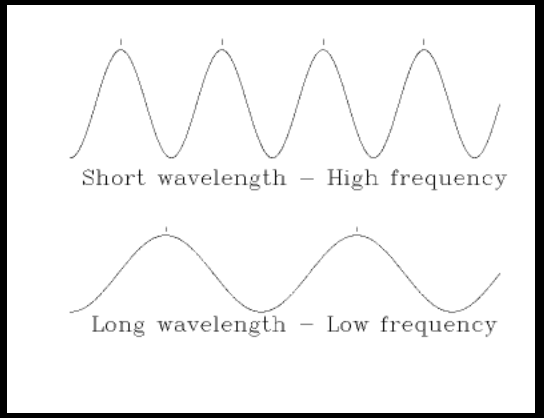
wavelength decreases, energy increases

frequency and energy

Wave energy & frequency are directly related.

frequency increases, energy increases

energy decreases, frequency decreases

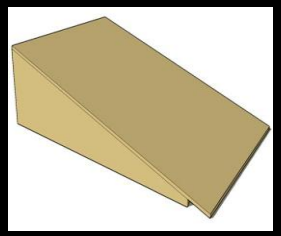


Differentiate between continuous & quantized

Continuous – white light spectrum

Quantized – emission spectra

Anything with "steps" is quantized



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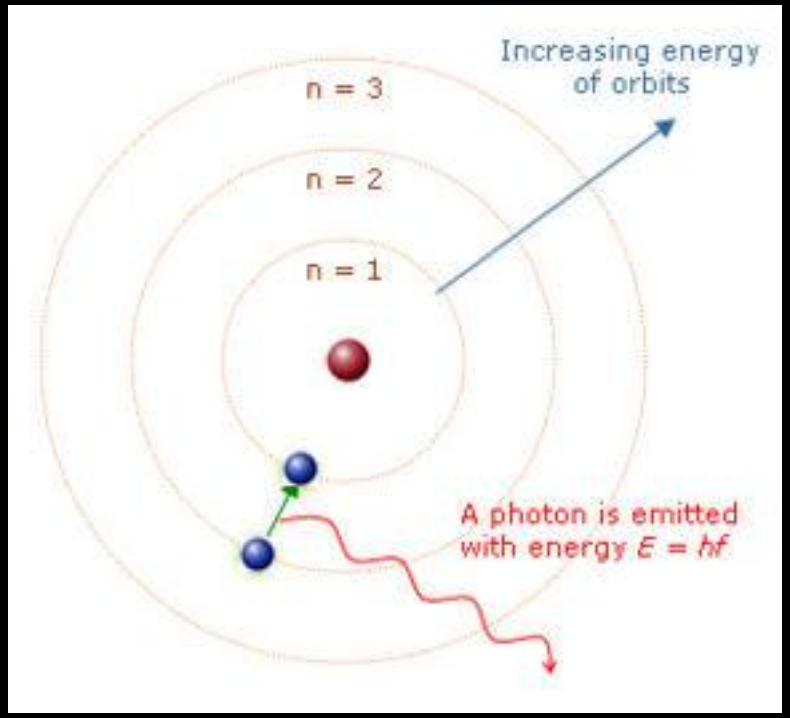
Distinguish between ground state & excited state of an atom

Ground state = lowest energy; resting state

Excited State = result of absorbing energy

moves electron to higher energy orbital



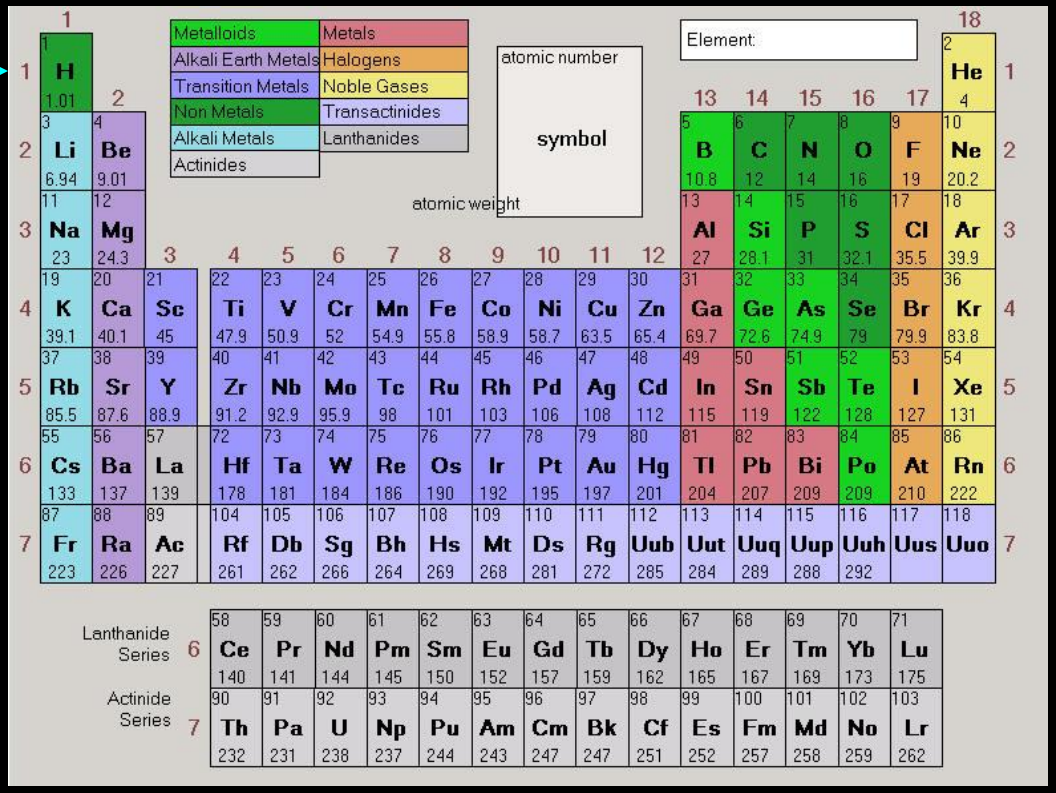
LPT

Identify the principal energy levels in an atom

Quantum number n

Corresponds to row of periodic table

n



State the energy trend among the principal energy levels in an atom

As n increases (move down rows of periodic table), energy increases

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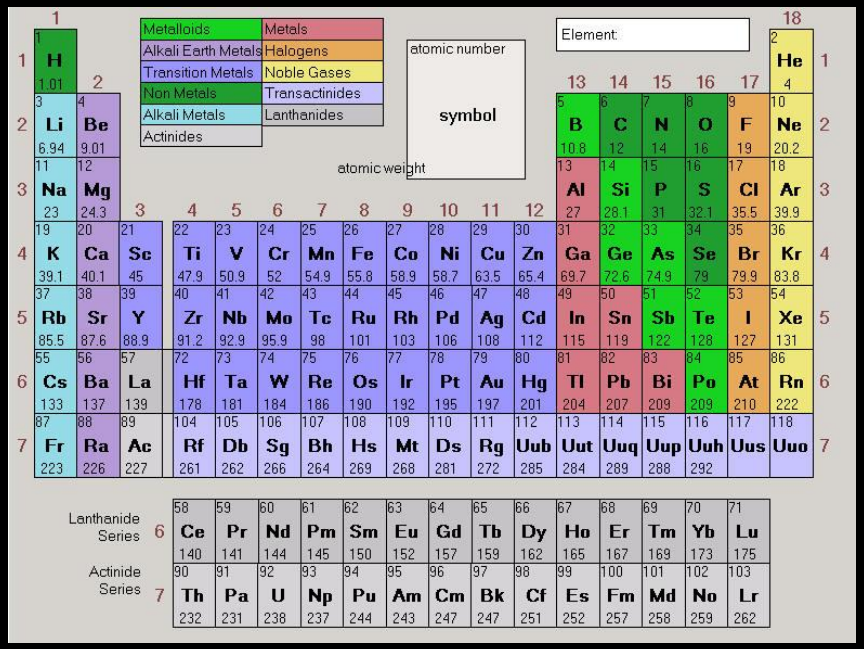
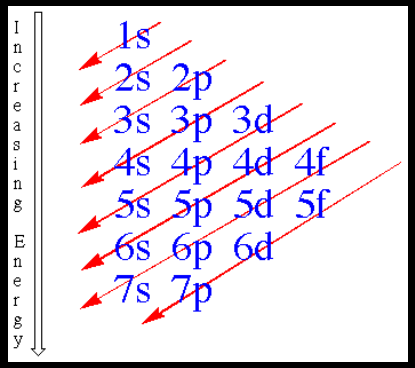
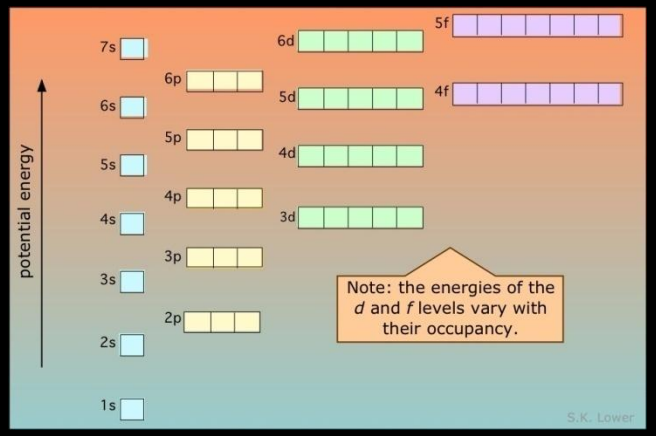
For each principal energy level, state the number of sublevels identify them.



n	Subshells (Orbitals)
1	s
2	s, 2p
3	s, 3p
4	s, 3d, 4p
5	s, 4d, 5p
6	s, 4f, 5d, 6p
7	s, 5f, 6d, 7p

State the relative energy trend among sublevels





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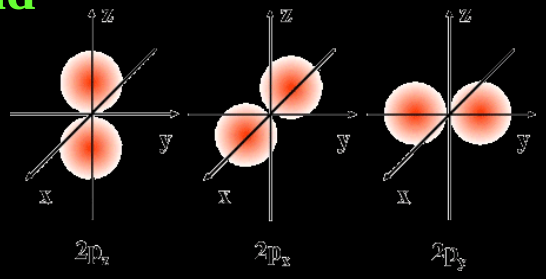
Describe what is meant by orbital

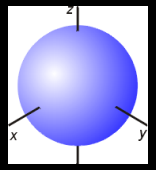
3-D plots of Ψ^2 functions represent electron orbitals

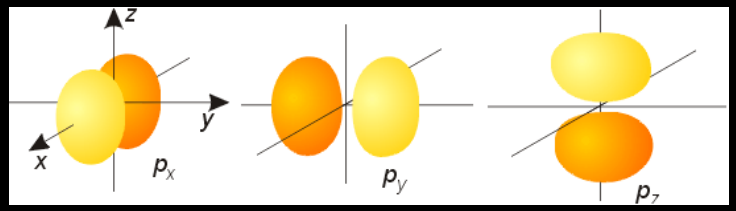
Region of space where an electron is likely to be found



Describe the shapes of s and p orbitals

s = spherical **p = dumb-bell shaped**





Recognize that the Pauli exclusion principle limits the number of electrons

Describe the restriction

Each sub-orbital can have a maximum of 2 electrons

(orbital may have 0, 1 or 2 electrons)

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Recognize chemical properties depend on electron configuration

Periodicity (Columns) a Function of Similar Outer Shell

Outer shell = valence electrons

Write ground state electron configuration for elements of Z = 1 – 36

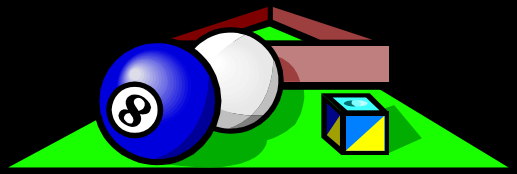
1	H	$1s^1$
2	He	$1s^2$ = [He]
3	Li	[He] $2s^1$
4	Be	[He] $2s^2$
5	B	[He] $2s^2\ 2p^1$
6	C	[He] $2s^2\ 2p^2$
7	N	[He] $2s^2\ 2p^3$
8	O	[He] $2s^2\ 2p^4$
9	F	[He] $2s^2\ 2p^5$
10	Ne	[He] $2s^2\ 2p^6$ = [Ne]
11	Na	[Ne] $3s^1$
12	Mg	[Ne] $3s^2$
13	Al	[Ne] $3s^2\ 3p^1$
14	Si	[Ne] $3s^2\ 3p^2$
15	P	[Ne] $3s^2\ 3p^3$
16	S	[Ne] $3s^2\ 3p^4$
17	Cl	[Ne] $3s^2\ 3p^5$
18	Ar	[Ne] $3s^2\ 3p^6$ = [Ar]
19	K	[Ar] $4s^1$
20	Ca	[Ar] $4s^2$
21	Sc	[Ar] $4s^2\ 3d^1$
22	Ti	[Ar] $4s^2\ 3d^2$
23	V	[Ar] $4s^2\ 3d^3$
24	Cr	[Ar] $4s^1\ 3d^5$*
25	Mn	[Ar] $4s^2\ 3d^5$
26	Fe	[Ar] $4s^2\ 3d^6$
27	Co	[Ar] $4s^2\ 3d^7$
28	Ni	[Ar] $4s^2\ 3d^8$
29	Cu	[Ar] $4s^1\ 3d^{10}$*
30	Zn	[Ar] $4s^2\ 3d^{10}$
31	Ga	[Ar] $4s^2\ 3d^{10}\ 4p^1$
32	Ge	[Ar] $4s^2\ 3d^{10}\ 4p^2$
33	As	[Ar] $4s^2\ 3d^{10}\ 4p^3$
34	Se	[Ar] $4s^2\ 3d^{10}\ 4p^4$
35	Br	[Ar] $4s^2\ 3d^{10}\ 4p^5$
36	Kr	[Ar] $4s^2\ 3d^{10}\ 4p^6$

LPT

Using n for the highest occupied energy level, write valence electron configurations of any representative element

Family	Outer Shell
Group1A	ns^1
Group 2A	ns^2
Group 3A	ns^2np^1
Group 4A	ns^2np^2
Group 5A	ns^2np^3
Group 6A	ns^2np^4
Group 7A	ns^2np^5
Group 8A	ns^2np^6

Octet Rule



Atoms want:
To be "noble"
To fill Outer shell to 8 valence electrons

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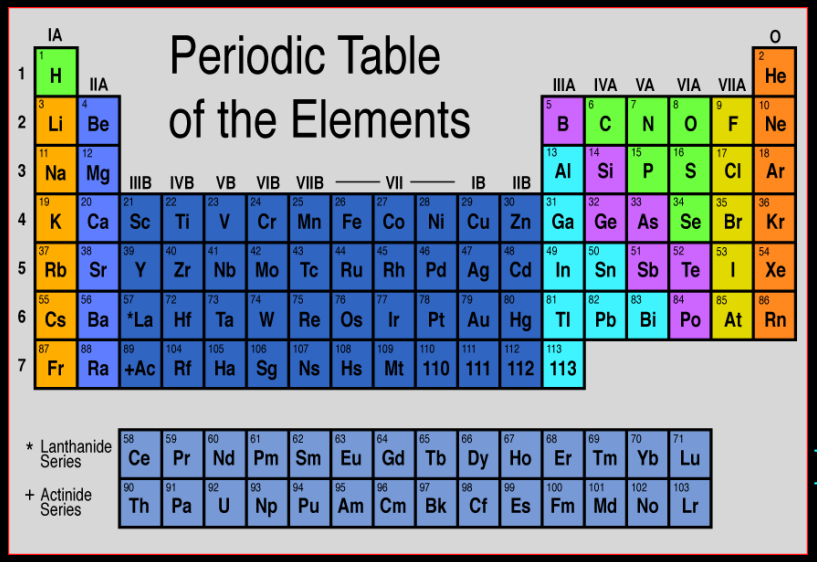
Write the Lewis (electron dot) symbol for an atom of any representative element

I	II	III	IV	V	VI	VII	0
H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca	Ga	Ge	As	Se	Br	Kr
Rb	Sr	In	Sn	Sb	Te	I	Xe
Cs	Ba	Tl	Pb	Bi	Po	At	Rn

LPT

Given symbol for a representative element, select other elements that would be expected to have similar chemical properties & conversely, elements that have different chemical properties.

Group 1A (1): alkali metals
Group 2A (2): alkaline earth metals
Group 7A (17): halogens
Group 8A (18): noble (inert) gases
Representative (1-2; 13-18): The A Groups (the Edges)
Transition Metals (3-12): The B Groups (the Center)
Metalloids: "Staircase" B,Si, Ge, As, Sb, Te, Po



Lanthanides = upper, of lower rows
Actinides = lowest, of lower rows

Predicted Chemical Properties
Elements in the same column are similar
Elements in different columns are different

LPT

Identify monatomic ions that are isoelectronic with a given noble gas; write the electron configuration of those ions.

Monatomic Ions With Noble Gas Electron Configurations

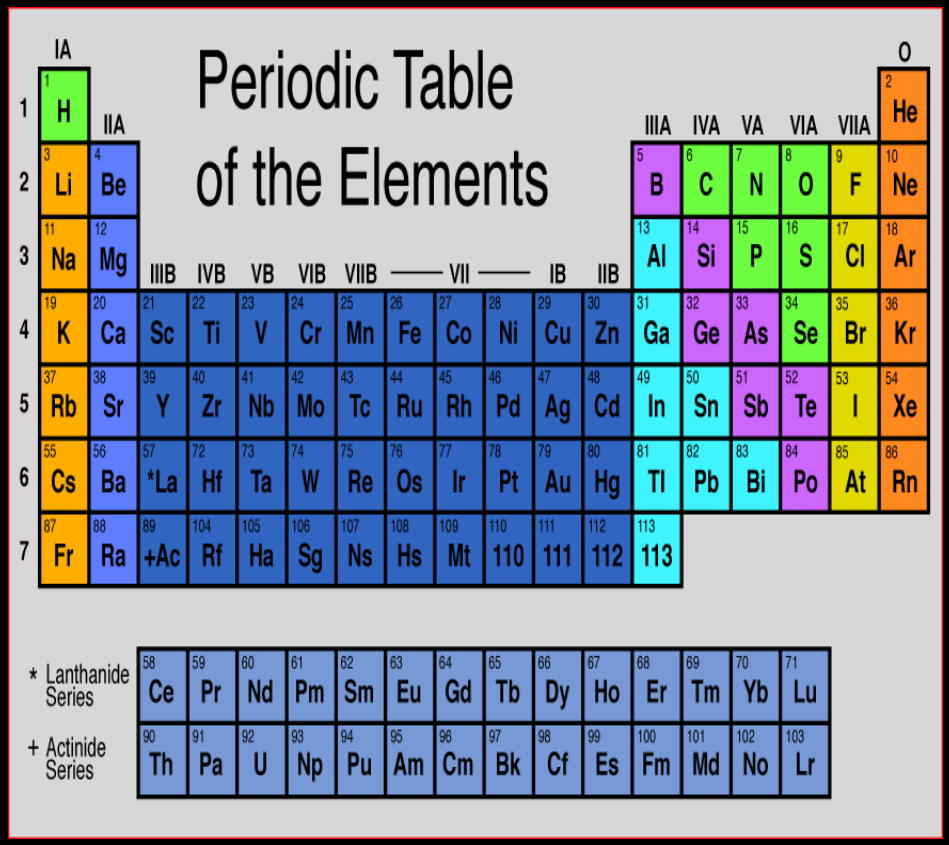
Isoelectronic = identical electron configuration

Atoms form ions to obtain a noble gas electron configuration

Na	$1s^2 2s^2 2p^6 3s^1$
Na^+	$1s^2 2s^2 2p^6$]
Ne	$1s^2 2s^2 2p^6$]

Isoelectronic

O	$1s^2 2s^2 2p^4$
O^{2-}	$1s^2 2s^2 2p^6$]
Ne	$1s^2 2s^2 2p^6$]



LPT

Distinguish between ionic and covalent bonds

Ionic:

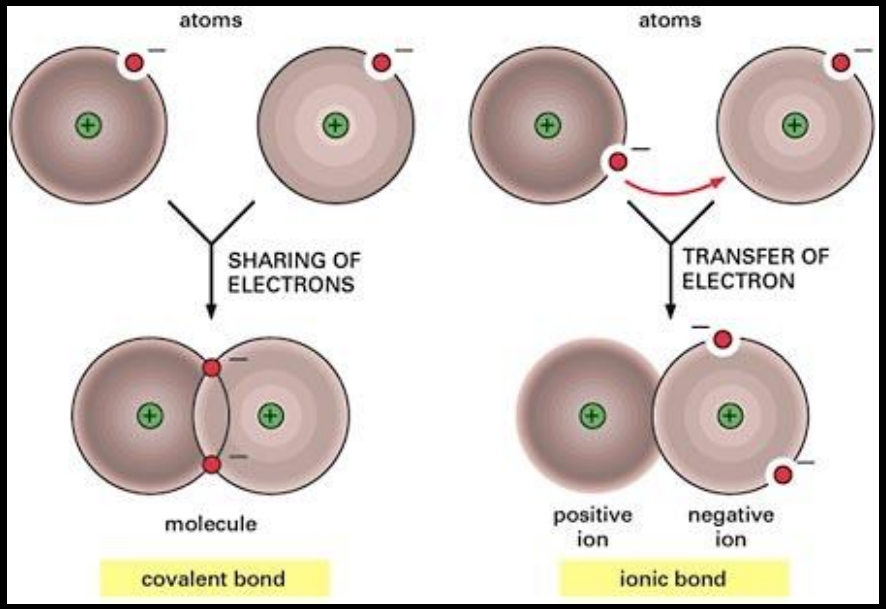
Transfer of electrons from one atom ➔ ions

+/- ions attracted to one another

Strong electrostatic forces hold ions within crystal matrix

Covalent:

Sharing a pair of electrons between two nuclei



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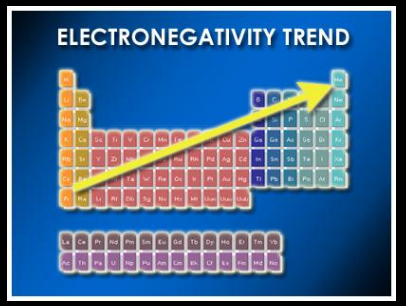
Differentiate between properties of ionic and covalent (molecular) compounds.

	Ionic	Covalent
Basic Component	Ions	Atoms/Molecules
Constituents	Metal + Non-Metal	2 Non-Metals
State (RT)	Solid	Solid, Liquid, Gas
Melting Point	Very High > 200 °C	Lower < 200 °C
Conductivity	Solids: Poor Melted: Good Aqueous: Good	Solids: Poor Melted: Good Aqueous: Good

LPT

Distinguish between polar and nonpolar covalent bonds

Nonpolar Covalent: equal sharing of e^- (Like C-H)
Polar Covalent: unequal sharing of e^- (Like O-H)



Given the electronegativities of all elements involved, rank bonds in order of increasing or decreasing polarity

$\Delta < 0.4$ ➔ non-polar covalent

Δ 0.4 - 1.9 ➔ polar covalent

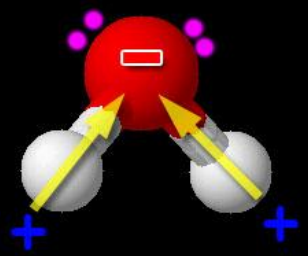
$\Delta > 1.9$ ➔ ionic

Δ = difference in electronegativity of the bonded atoms

>difference, greater the polarity of the bond

If the bond is polar, state which end is positive and which end is negative.

Most electronegative atom ➔ negative end of dipole



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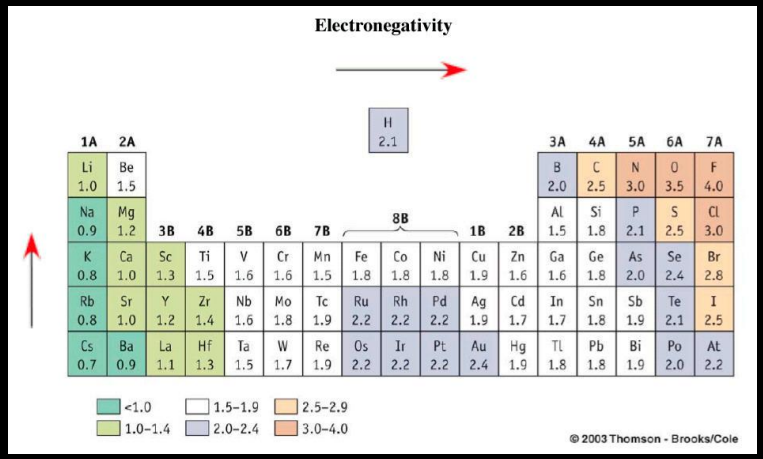
Using an electronegativity table, determine the bond type for the following:

H-Cl (3.0 – 2.1 = 0.9) = polar covalent
Ca-Cl (3.0 – 1.0 = 2.0) = ionic
P-Cl (3.0 – 2.1 = 0.9) = polar covalent
Se-Br (2.8 – 2.4 = 0.4_ = non=polar covalent
O-F (4.0 – 3.5 = 0.5) = polar covalent
O-H (3.5 – 2.1 = 1.4) = polar covalent



At which atom is the negative end of the dipole?

O-H (O = 3.5; H = 2.1) O ➔ most electronegative is negative (arrow) end
Si-Cl (Cl = 3.0; Si = 1.8) Cl

Which is the most polar bond … O-H or SH?

O-H (3.5 – 2.1 = 1.4) ➔ greater electronegativity difference
S-H (2.5 – 2.1 = 0.4)

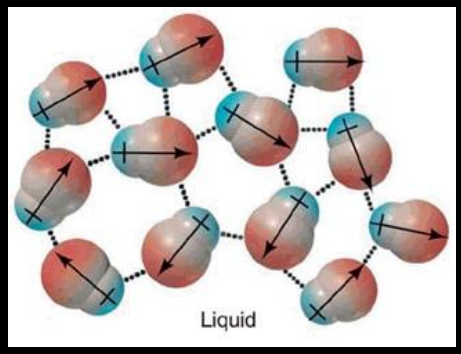
LPT

Identify and describe or explain dipole forces, dispersion forces, & H- bonds

Dipole-Dipole Interactions:

Molecules with permanent Dipole

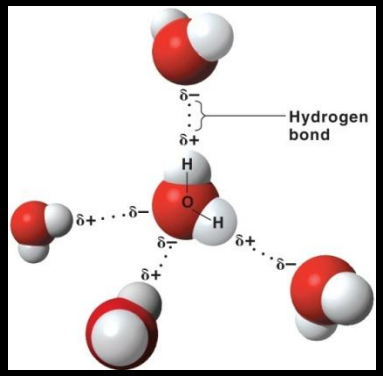
Dipoles align … cohesive attraction



Hydrogen Bonds

H & electronegative atom (especially N & O; F)

Very important In biological systems



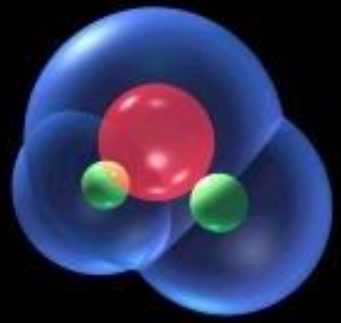
London Dispersion (Van derWaal's) Forces

Weakest interaction

Temporary; when adjacent atom electrons create dipole

All atoms; more prevalent in heavier/larger

Stronger when atoms easily polarized



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Given the structure of a molecule, or information from which it may be determined, identify the significant intermolecular forces present.

Dipole-Dipole Interactions require polar bonds/atoms

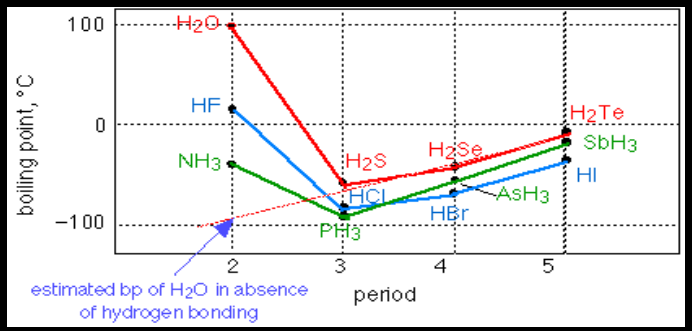
Hydrogen bonds require H and electronegative atom

London Dispersion – available to all atoms

Given the molecular structure of two substances, or info from which they may be obtained, compare or predict relative values of physical properties

If only dispersion forces present (no H-bonding), the more electrons present (higher Z), > boiling point

If H-bonding (> dispersion forces) present, H bonded higher, Then, the more electrons present (higher Z), > boiling point



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