Big Idea #2

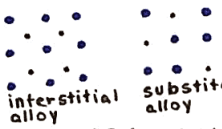
Bonding & Phases

ionic bonds occurs between metals and nonmetals, cation "gives up" electron to anion
melting points very high (smaller ions → less coulombic attraction → higher melting point)
boiling points very high
conductivity poor in solids (electrons don't move in lattice), liquids can be conductive
generally solid @ room temp

greater ion charges → stronger lattice energy (based on Coulomb's Law)

smaller ions → stronger coulombic attractions between nuclei

metallic bonds
conductivity high
malleable, ductile
occurs between metals



covalent bonds atoms share electrons to complete valence shells

bond type	single	double	triple
bond designation	one σ	one σ, one π	one σ, two π
bond order	1	2	3
bond length	long	intermediate	short
bond energy	least	intermediate	greatest

network covalent bonds
conductivity poor
melting point very high
boiling point very high

p-doping atoms added with fewer e^-, creates positive charge
n-doping atoms added with extra e^-, creates negative charge
doping increases conductivity by compromising lattice

resonance forms
total bond order = $\frac{\text{sum of bond orders across resonance forms}}{\text{number of resonance forms}}$

$\frac{\text{number of bonds}}{\text{number of atoms}}$ = bond order

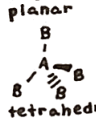
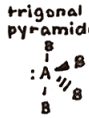
incomplete octets
hydrogen
helium
boron

expanded octets
atoms from n=3 with empty d-orbitals
ex) P, S, Xe

formal charge
number of valence electrons for atom − number assigned for Lewis structure
↳ lone pairs = 2 ve^-
↳ bonds = 1 ve^-

fewer formal charges → more likely Lewis structure
or negative charge on more electronegative atom

molecular geometry

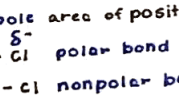
e^- pairs	hybridization	basic shape	other geometries
2	sp 180°	B − A − B linear	
3	sp^2 120°	 trigonal planar	 bent
4	sp^3 109°	 tetrahedral	 trigonal pyramidal  bent
5	sp^3d^1 120° and 90°	 trigonal bipyramidal	 seesaw  t-shaped  linear
6	sp^3d^2 90°	 octahedral	 square pyramidal  square planar

multiple bonds take up more space than single bonds
lone pairs reduce the bond angles between terminal angles

polarity electrons are unequally shared in a bond
dipole area of positive or negative charge
$\delta+ \; \delta-$
C − Cl polar bond
Cl − Cl nonpolar bond

molecular polarity
symmetrical structures tend to be non-polar

intermolecular forces
dipole-dipole occur between polar molecules (tend to be pretty weak)
hydrogen bonds occur between hydrogen end of molecule and F, O, N of another molecule



london dispersion forces very weak attractions between all molecules
more electrons → stronger LDFs (more polarizable)

bond strength
ionic substances − based of coulombic attractions
covalent substances − hydrogen > dipole > LDF
lower melting / boiling point than ionic
metallic bonding − very strong
network covalent − strongest type of bond

vapor pressure
stronger IMFs → lower vapor pressure

solution separation
more ions from dissociation → higher conductivity
solute substance being dissolved
solvent substance doing the dissolving

paper chromatography

$R_f = \frac{\text{distance travelled by solute}}{\text{distance travelled by solvent}}$
(retention factor)

column chromatography
analyte molecules attracted to eluent
↳ faster analyte = more polar

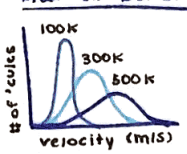
distillation
uses differing boiling points to separate / collect substance

deviations from ideal behavior
significant gas volume → larger than predicted volume
gas molecules attracted to each other → real pressure smaller than predicted

kinetic molecular theory
$KE = \frac{1}{2}mv^2$
m = mass of molecule (kg)
v = speed of molecule (m/s)
KE measured in J

all gases have same KE at same temps
assumes insignificant volumes, no attraction between gas molecules

Maxwell-Boltzmann Diagrams



effusion rate at which gas escapes through microscopic holes
faster molecules → higher effusion rate

Dalton's Law
$P_{total} = P_a + P_b + P_c \ldots$
$P_a = (P_{total})(X_a)$
X_a = moles of gas A / total moles of gas

ideal gas equation
$PV = nRT$
P = pressure of gas (atm)
V = volume of gas (L)
n = moles of gas
R = 0.0821 L·atm/mol·K
T = temp of gas (K)

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
pressure increases → temp increases
pressure increases → volume decreases (Boyle's Law)
volume increases → temp increases (Charles's Law)