AP Chemistry Midterm Study Guide

Chapter 1: Matter and Measurement

- Chemistry: the science that seeks to understand the behavior of matter by studying the behavior of atoms and molecules
- Scientific method:
    - Hypothesis: a tentative interpretation or explanation of observations; should be falsifiable
    - Experiment: a controlled procedure to generate observations to support or refute a hypothesis
    - Scientific law: a statement that explains past observations that can be used to predict future ones
        - e.g., law of conservation of mass: matter neither created nor destroyed in chemical reaction
    - Scientific theory: a model to help explain nature and made up of multiple well-established hypotheses, laws, and/or other theories
        - e.g., atomic theory: explains law of conservation of mass and proposed that atoms were small, indestructible particles
- Classification of matter:
    - Matter: something that occupies space and has mass
    - Substance: a specific instance of matter, classified by state and composition
        - Pure substance: substance made of only one particle, fixed composition
    - Mixture: matter composed of different types of particles, varying composition
        - Heterogeneous mixture: when the composition varies from one section to another; components can be physically separated
        - Homogeneous mixture: when the composition is uniform throughout
            - Solution: a liquid homogeneous solution
        - Methods of separating mixtures
            - Decanting: method of separation of heterogeneous solution of an insoluble solid and a liquid (pouring the liquid off)
            - Filtration: method of separation of heterogeneous solutions of an insoluble solid and a liquid (sieving the liquid through)
            - Distillation: method of separation of homogeneous solutions by utilizing different boiling points
    - Element: a pure substance that cannot be broken down smaller
        - Atoms: the building blocks of matter, the smallest particle that has all the properties of that element
    - Compound: a substance composed of multiple atoms joined together chemically
        - Molecules: the combination of atoms by covalent bonds
        - Formula units: the unit that expresses the ratio of atoms in an ionic compound
- States of matter:
    - Solid: atoms vibrate but do not shift past one another; fixed volume and shape

        - Amorphous (unstructured) or crystalline (structured in repeating pattern)
    - Liquid: atoms move past one another freely but are packed together; fixed volume but not shape
    - Gas: atoms move past one another freely and occupy all free space; free volume and shape
- Physical and chemical changes:
    - Physical changes: changes altering state or composition of a substance, but not composition
    - Chemical changes: changes that alter composition of a substance, usually through a chemical reaction
- Physical and chemical properties:
    - Physical property: a property that is displayed without changing a substance's molecular composition
        - Can be intensive (not based on amount of substance, e.g., density) or extensive (based on amount of substance, e.g., mass) properties
    - Chemical property: an intensive property that is displayed only when changing a substance's molecular composition (i.e., an observation relating to a chemical change)
- Volatility: property of a substance indicating its ability to boil (easy to boil = low boiling point = volatile)
- Energy: the capacity to do work
    - Work: the action of a force through a distance
    - Energy of a substance is sum of kinetic energy (energy of movement) and potential energy (stored energy associated with position or composition) associated with object
        - Thermal energy and movement are examples of kinetic energy
        - Chemical potential energy and position are examples of potential energy
    - Law of conservation of energy states that energy can change from one form into another, but cannot be lost— total amount in the universe is constant
    - Systems tend to move from high energy to low energy; high-energy systems are unstable
- International Standard of Units (SI) system based on metric system (as opposed to English system)

- Unit multiplier prefixes:

Prefix	Symbol	Multiplier	
exa	E	10^{18}	1,000,000,000,000,000,000
peta	P	10^{15}	1,000,000,000,000,000
tera	T	10^{12}	1,000,000,000,000
giga	G	10^{9}	1,000,000,000
mega	M	10^{6}	1,000,000
kilo	k	10^{3}	1,000
hecto	h	10^{2}	100
deka	da	10^{1}	10
deci	d	10^{-1}	0.1
centi	c	10^{-2}	0.01
milli	m	10^{-3}	0.001
micro	μ	10^{-6}	0.000,001
nano	n	10^{-9}	0.000,000,001
pico micromicro	p $\mu\mu$	10^{-12}	0.000,000,000,001
femto	f	10^{-15}	0.000,000,000,000,001
atto	a	10^{-18}	0.000,000,000,000,000,001

- Significant figures are a system of preserving precision of data through measurements and calculations
  - They are measured so that every digit is estimated, except for the last (measurement plus guessed digit)
  - Exact (counting) numbers have infinite sigfigs, never lose precision
  - Multiply / divide → answer has # sigfigs of the multicand/multiplier/dividend/divisor with least sigfigs
  - Add / subtract → answer has precision to multicand/multiplier/dividend/divisor with least precision
- Precision "refers to how close a series of measurements are to one another or how reproducible they are"
  - Systematic error gives biased probability of too high or too low; higher precision but lower accuracy
- Accuracy "refers to how close the measured value is to the actual value"
  - Random error gives equal probability of too high or low; higher accuracy but lower precision
- Dimensional analysis is a method of converting between units by using conversion factors

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Chapter 2: Atoms, Molecules, and Ions

- Scanning tunneling microscopy (STM) allows for the imaging of and moving of individual atoms and molecules
- Law of conservation of mass states that matter is neither created nor destroyed in a chemical reaction
- The atomic model developed through multiple scientists
  - Democritus (ancient Greece) thought that atoms were indestructible, tiny particles

    - Dalton built upon the recent ideas of other scientists (such as Copernicus, Kepler, and Newton) to form a more modern atomic theory
- Law of definite proportions states that "all samples of a given compound, regardless of their source or how they were prepared, have the same proportions of their constituent elements"; ratio of elements in compound always the same
- Law of multiple proportions states that "when two elements (call them A and B) form two different compounds, the masses of element B that combine with 1g of element A can be expressed as a ratio of small whole numbers"
- J.J. Thomson discovered the electron, a negatively-charged subatomic particle
    - Cathode rays were shot from one side to another in a cathode ray tube and composed of electron particles
    - Electrons have electrical charge, which create an electric field
- Robert Millikan discovered the charge of an electron using the oil drop experiment, which suspended oil drops based on their charge and size
- Rutherford discovered the nucleus and the proton, which formed the dense, positively-charged center of an atom
    - Chadwick discovered that there were also neutrons in the nucleus
- The atomic mass unit (amu) is a common unit for measuring the mass of subatomic particles
    - 1/12 of the weight of a C-12 atom
- Elements are characterized by atomic number (Z) (number of protons in nucleus) and chemical symbol (1-3 letter representation, first letter capitalized)
    - Elements can vary in number of neutrons; isotopes have same # protons but varying neutrons
        - Natural abundance is the percentage of a certain isotope in a natural sample of an element
        - Sum of neutrons and protons = mass number (A) (approximately equal to atomic mass)
            - Isotopes can be represented like so: $^{12}_{6}C$ (for carbon with six neutrons) (mass number superscript, atomic number subscript, chemical symbol on right)
            - Isotopes can be represented like so: carbon-12 or C-12 (for carbon with six neutrons) (element name or formula and then mass number)
        - Isotopes only vary in nuclear stability; chemical and physical properties do not change (dependent on valence electrons and bonding)
    - Elements can vary in number of ions; ions have differing number of electrons and are charged accordingly
        - Positive ions (less electrons than protons) are called cations
        - Negative ions (more electrons than protons) are called anions
- Dmitri Mendeleev ordered the elements according to the periodic law, which states that elements arranged by increasing mass would have recurring (periodic) chemical patterns
    - Moseley later changed periodic law to order by increasing periodic number
- Atomic mass is the weighted average mass of the isotopes of an element based on their relative abundances

- Mass spectrometry involves the ionization of atoms and then shooting them in a trajectory altered by a magnet into a detector; the heavier particles are affected least and vice versa for the lighter particles
- A mole (mol) of a substance is Avogadro's number of it
    - An atom's molar mass is numerically equal in grams per mole as amu per atom (atomic mass)

Chapter 3: Mass Relations in Chemistry: Stoichiometry

- Compounds are very different than their component pieces
- Empirical formula is a chemical formula with the lowest ratios of elements
    - Always the case for formula units (ions)
- Molecular formula is a chemical formula that actually occurs for molecules
    - May be the same as the empirical formula
- Structural formula shows bond angles (similar to Lewis structure)
- Ball-and-stick model shows bonding of particle and shows shape
- Space-filling model shows approximate proportions of atoms and how it would look in real life
- Atomic elements (e.g., Na) vs. molecular elements (e.g., BrINClHOF diatomics or S_6)
- Common names (e.g., water) vs. systematic names (e.g., dihydrogen monoxide (???))
- Mass percent composition is percentage of mass of element to mass of compound
- Combustion analysis is used to calculate empirical formulas of substances
    - $H \rightarrow H_2O$
    - $C \rightarrow CO_2$
    - $O \rightarrow H_2O$ and CO_2 (subtract sum of masses of Hs and Cs from original)
- Organic compounds mostly carbon and hydrogen (especially as hydrocarbons), sometimes N, O, S
    - Carbon always forms four bonds (i.e., not substances like CO_2 w/ double bonds)
    - Hydrocarbons can have functional groups such as hydroxyl (-OH) groups
        - Hydrocarbons with same functional group form a family

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Chapter 4: Chemical Quantities and Aqueous Solutions

- Reaction stoichiometry
- Limiting reactant, theoretical yield, and percent yield
- Solution concentration, molarity
    - Stock solutions are highly-concentrated solutions from which to make laboratory solutions of different concentrations
        - Use $M_1V_1 = M_2V_2$ to convert to lower molarity solution
- Electrolytes are substances that dissolve in water and conduct electricity as an aqueous solution
    - The better they dissolve, the stronger they are and the better the electricity they conduct
    - Often ionic compounds are good electrolytes
- Precipitate reactions happen when two solutes react to form a precipitate (an insoluble solid)
    - Molecular equations are the full equation

    - Complete ionic equations have each aqueous compound as ions
    - Net ionic equations have reacting elements only (i.e., without spectator ions)
- Acid-base and gas-evolution reactions
- Oxidation-reduction reactions
- *Lab*: Gravimetric Analysis of Calcium and Hard Water

Chapter 5: Gases

- Measurement of gases
- The ideal gas law
    - Boyle's Law (P-V), Charles' Law (V-T), Guy-Lussac's Law (P-T), Avogadro's Law (n-V)
- Gas law calculations
    - Dalton's Law (sum of partial pressures = total pressure)
- Mixtures of gases and partial pressures
- Stoichiometry of gaseous solution
- Kinetic Molecular Theory (KMT) of gases
    - Volume of gas particle is 0
    - Average kinetic energy is proportional to temperature in kelvins
        - Kinetic energy = $\frac{1}{2}mv^2$
    - Collision of particles is elastic
- Mean free path, diffusion, and effusion of gases
    - Use root mean square velocity for velocity of gases
    - Mean free path decreases with pressure and particle size
    - Graham's law of effusion
- Real gases: the effects of size and intermolecular forces
    - At high pressures, volume is greater because of particle size and vice versa
    - At low temperatures, pressure is lower because of high intermolecular forces and vice versa
- *Lab*: Molar Volume of a Gas
- *Interactive*: PhET Virtual Lab-Gas Laws

Chapter 6: Thermochemistry

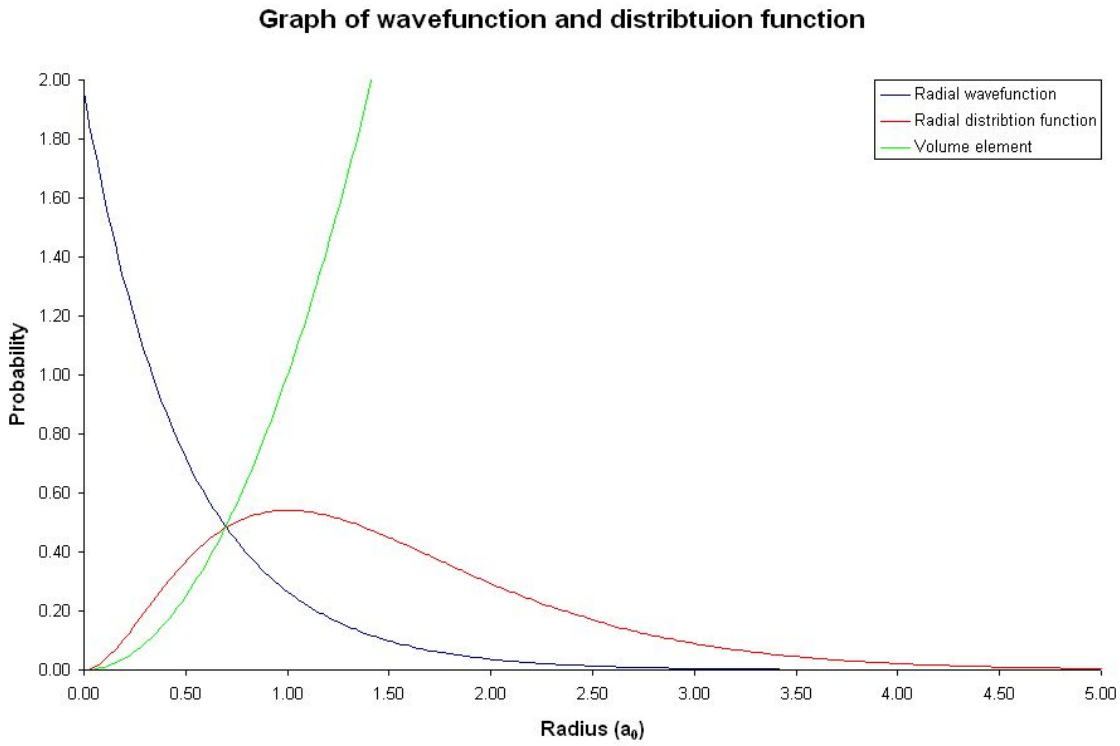
- First law of thermodynamics
- Heat, work, pressure-volume work
    - <u>Heat</u>: flow of energy through a distance
        - Transfer of energy between <u>system</u> and <u>surroundings</u> until it reaches thermal equilibrium
    - <u>Internal energy</u> (E) = KE + PE
        - E is a state function
    - <u>Pressure-volume work</u>: work is done when pressure or volume increases
        - There must be a change in temperature for this to happen
- Calorimetry and heat

- Enthalpies
    - Enthalpy (H) = E + PV
        - $\Delta H = q$ (change in enthalpy = heat, pressure-volume work is cancelled out)
- Thermochemical equations
    - Hess's Law: adding and subtracting thermochemical equations and enthalpies of reactions
    - Standard enthalpy of formation: heat of forming one mole of compound at standard state (kJ/mol)
    - Standard state: 1atm and 25°C
- Relationship between heat and other forms of energy—measurement of first law of thermodynamics
- *Lab*: Designing a Hand Warmer
- *Lab*: Enthalpy of Reaction and Hess's Law

Chapter 7: The Quantum-Mechanical Model of the Atom

- Schrödinger's cat
    - Presents the quantum-mechanical model of the atom, which presents light and particles as both particles and waves
- Light, photon energies, and atomic spectra
    - Light is electromagnetic radiation, with certain properties:
        - Amplitude (intensity)
        - Wavelength (ƛ) = c/v (color)
        - Frequency (v) = c/ƛ (color)
        - Speed (c) = vƛ
    - Electromagnetic spectrum contains all wavelengths/frequencies of electromagnetic radiation
        - Energy comparison (greater to lesser)
            - Gamma ray
            - X-ray
            - Ultraviolet ray
            - Visible light
            - Infrared waves
            - Microwaves
            - Radio waves
    - Photoelectric effect shows the particle nature of light
        - Light is made up of photons, bundles of light energy
        - Photons are quantized, with distinct amounts of energy
            - E = hv
- Atomic orbitals: shapes and sizes
- PES (photoelectron spectroscopy)
    - Atomic spectroscopy: study of electromagnetic radiation absorbed and emitted by atos
    - Emission spectrum is unique to every element

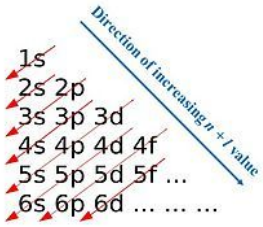
        - Bohr created a model of the atom with quantized energy levels
            - Quantization of electrons shows wave nature
            - $\Delta E = -2.18 \times 10^{-18}(\frac{1}{n_f^2} - \frac{1}{n_i^2})$
- de Broglie equation for the wavelength of particles
    - $\lambda = \frac{h}{mv}$
- Uncertainty principle: never observe both wave and particle natures of matter or light
    - This is the nature of the particle-wave duality
    - Particle nature and wave nature are complementary properties
- Quantum numbers: notation to indicate an electron's position in an atom
    - Principle (n): energy level (1 to 7)
    - Angular momentum (l): sublevel (0 to (n-1))
    - Magnetic (m_l): orbital position within sublevel (-n to n)
    - Spin (m_s): electron spin (±½)
- Orbital: probability map for up to two electrons, defined by Schrodinger's wave function
- Probability density: probability per unit volume of finding electron in space
- Radial distribution function (total radial probability): represents total probability of finding electron within spherical shell at distance r from nucleus



- *Lab*: Discharge Lamps and Flame Tests
    - See atomic spectroscopy / photoelectron spectroscopy
- *Lab*: PES
    - See atomic spectroscopy / photoelectron spectroscopy

Chapter 8: Periodic Properties of the Elements

- Periodic table history
- Electron configurations and orbital notation
    - Electron configuration organized by aufbau principle (increasing energies)

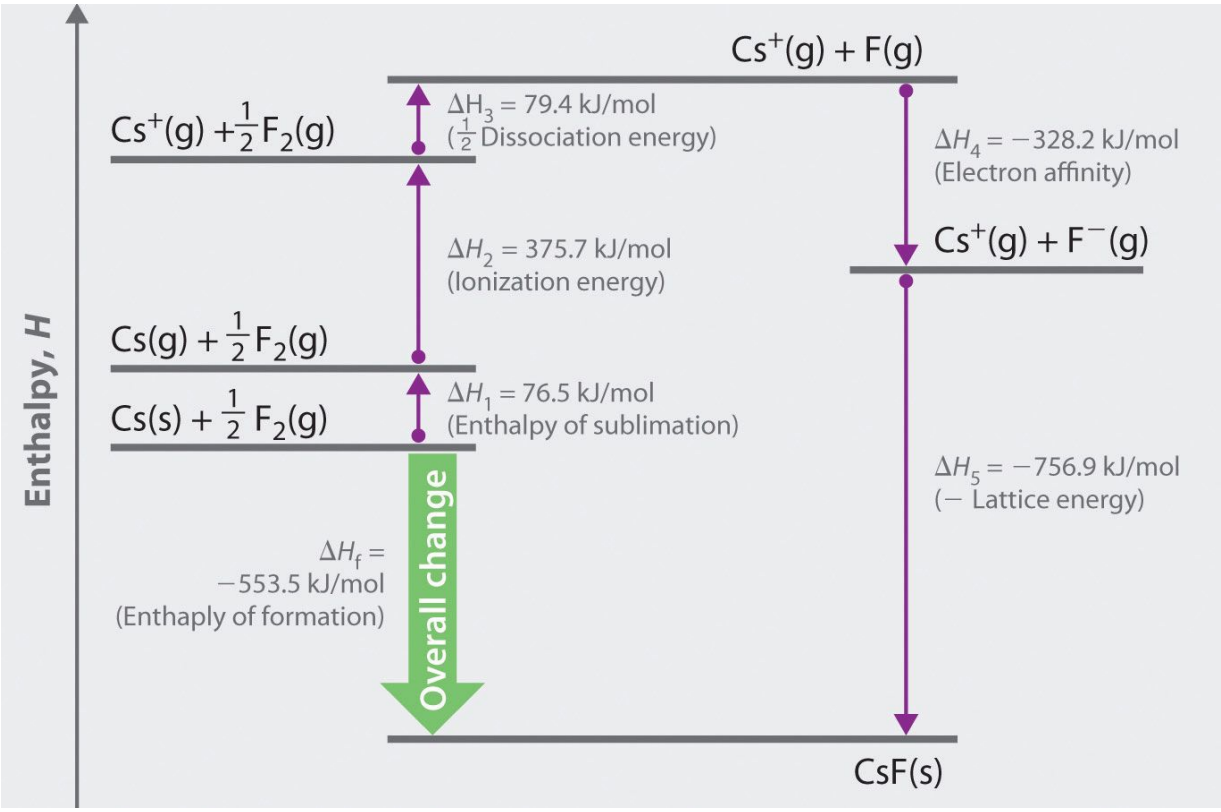


    - Pauli exclusion principle: no two electrons can have the same four quantum numbers; two electrons in same orbital must have opposite spins
    - Hund's rule: orbitals in the same sublevel must fill themselves singly before filling doubly
    - Energy levels in H are degenerate (have same energy); in other atoms, sublevels are degenerate
- Periodic trends: size and effective charge
    - Coulomb's law: $E = \frac{kq_1q_2}{d^2}$
        - Directly proportional to charge and inversely proportional to distance
        - Helps explain most of the periodic trends
    - Electron shielding leads to effective nuclear charge
        - ENC = total nuclear charge + charge of shielded electrons
        - Penetration decreases shielding
    - Nonbinding atomic radius (Van der Waals radius): half distance between two adjacent atoms
    - Binding atomic radius (covalent radius): half distance between two bonded atoms
- Ions: ionic radius and ionization energy
    - Electrons are lost based on distance from nucleus (energy level), not necessarily reverse aufbau
    - Atoms are either paramagnetic (odd # electrons) or diamagnetic (even # electrons); ions usually diamagnetic (with stable, even # of electrons)
    - Exceptions in ionization energy when moving from s to p sublevels (p is further away, easier to take away even though it has more protons and electrons)
    - Electron affinity: energy change associated with gain of electron
- Metallic character

Chapter 9: Chemical Bonding I: The Lewis Model

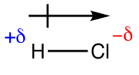
- Bonding theories are different chemical models of atoms
    - Lewis structures are a bonding theory that represents molecules with dots, dashes, chemical symbols, and brackets (for ions)

    - Valence bond theory is another bonding theory that treats an atom more in its quantum-mechanical state
    - Molecular orbital theory is the third bonding theory that treats an atom and its electrons in the quantum-mechanical model
- Chemical bonds form because they lower the energy of a molecule and thus are more stable
    - A chemical bond is "the sharing or transfer of electrons to attain stable electron configurations for the bonding atoms" (according to the Lewis model)
- Ionic bonding and lattice energies
    - Ionic bond is when an atom is ionized by another and the two atoms attract one another according to Coulomb's law
    - Ions can be represented in the Lewis model by brackets with the charge indicated, like so: $[Na]^+ [:\ddot{Cl}:]^-$
    - Lattice energy is the energy associated with the formation of the crystal lattice structure of an ionic compound
        - The creation of an ionic compound is always exothermic because it releases a lot of energy and becomes very stable as a crystal lattice structure because of the many nondirectional charges
        - The stronger the charges or the smaller the ionic radii, the stronger the lattice energy (according to Coulomb's Law)
        - High lattice energy of ionic compounds accounts for ionic compounds' high melting point, inability to conduct electricity as solids, and tendency to conduct electricity in solution
        - The Born-Haber Cycle can be used to calculate this
            - Heat of formation of the gaseous metal + the heat of formation of a single anion (if applicable, from diatomic molecule) + the ionization energy of the cation + the electron affinity of the anion + the lattice energy of the ionic compound = the heat of creation of the ionic compound



^ The Born-Haber Cycle for CsF (used for calculating lattice energy)^

- Covalent bonding: Lewis structures
  - Covalent bond is when two atoms *share* electrons
    - A shared pair of electrons in a covalent compound is called a bonding pair
      - A bonding pair can be represented as two dots or a line
    - An unshared pair of electrons in a molecule is called a lone pair
    - A double-bond is when two atoms share four electrons; a triple-bond is when two atoms share six electrons; in both cases, the electrons count towards the octets of both atoms
      - The greater number of bonds, the stronger (more bonding pairs = more charge = higher Coulombic attraction) and shorter (also according to Coulomb's law) the bond
    - Bond attractions are directional (in contrast to nondirectional ionic bonds)
      - This means that there is less attraction and easier bonds to break between molecules, leading to low boiling points
  - The Lewis model consists of Lewis (electron dot) structures, which have atoms represented by their chemical symbol and with their valence electrons represented as dots
    - Examples: chlorine: ·Cl· ; oxygen: ·O·
    - It has a high predictive power as to what is and is not likely to be a compound

        - Eight valence electrons forms an octet, which is usually very stable, hence the octet rule
            - Octet rule has exceptions with odd numbers of electrons (free radicals), incomplete octets (such as BF_3 or H_2), and expanded octets (such as SO_3)
- Metallic bonding and the sea of electrons
    - Metallic bond forms between atoms of a metal and happen when the metals give up electrons to become cations and the electrons hold the metal together
- Electronegativity and bond polarity
    - Electronegativity is the "ability of an atom to attract electrons to itself in a chemical bond"
    - A higher electronegativity in one atom in a covalent bond than another can leads to a polar covalent bond
        - Polar covalent bonds lie between pure covalent and ionic bonds
        - H-Cl is a polar covalent bond: 

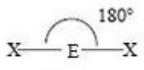
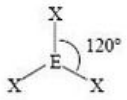
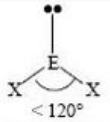
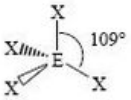
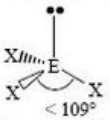
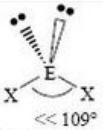
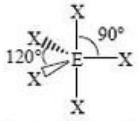
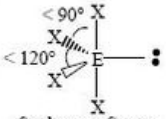
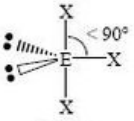
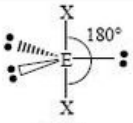
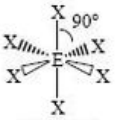
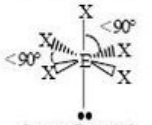
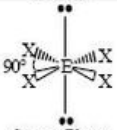
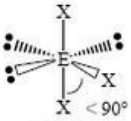
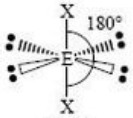
; the arrow points towards the more electronegative, and the delta-positive/delta-negative can also indicate the more electropositive or electronegative of the two, respectively
    - Electronegativity increases moving up and right in the periodic table
    - Dipole moment is when there is a difference in positive and negative charge
    - Percent ionic character is the "ratio of a bond's actual dipole moment to the dipole moment it would have if the electron were completely transferred from one atom to the other"
        - Over 50% percent ionic character (electronegativity difference > 1.7) means ionic
        - Less than 50% percent ionic character (electronegativity difference < 1.7) means covalent
- Resonance and formal charge
    - Resonance happens when more than one valid Lewis diagram can be drawn for a molecule
        - A resonance hybrid is an average, actual structure of a resonant molecule
    - Formal charge is "the charge [an atom] would have if all bonding electrons were shared equally between the bonded atoms"
        - FC = VE - ½BE - LE
        - Distinguishes between multiple valid Lewis structures or resonance structures to find the most likely one
        - The lower the formal charge, the better; on an ion, formal charges should sum to the overall charge of the ion; if charges are necessary, the most electronegative atom should hold the negative charge(s)
- Polarity of molecules
- *Lab*: What's in a Bottle?

Chapter 10: Chemical Bonding II: Molecular Shapes, VSEPR, and Molecular Orbital Theory

- VSEPR theory: the five basic shapes

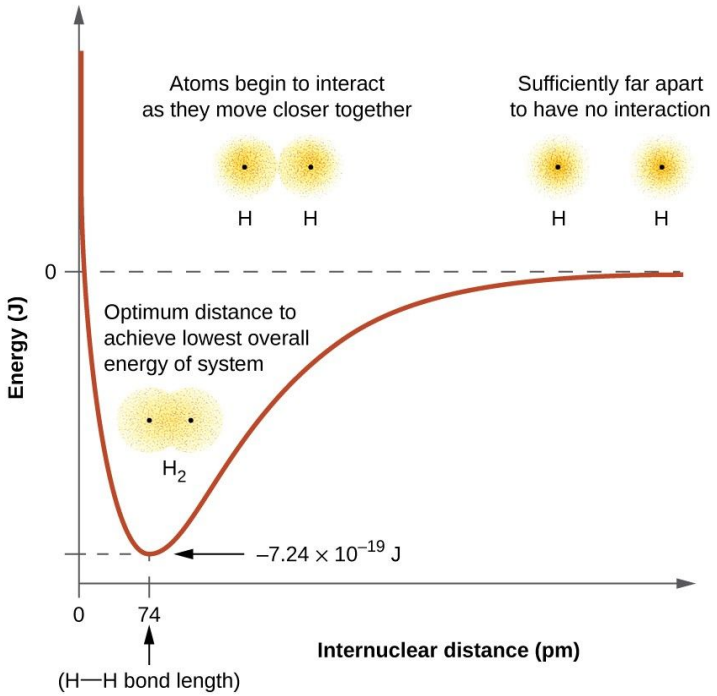
- Valence Shell Electron Pair Repulsion (VSEPR) theory is based off the idea that electron groups repel themselves through coulombic forces
  - Lone pairs occupy more space (more repulsion) than bonding pair
- Electron geometry based on all electron groups; molecular geometry based on bonded groups
  - They are the same when all electrons of atom involved in bonds (o lone pairs)

VSEPR Geometries

Steric No.	Basic Geometry 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

- Lone pairs
- Molecular geometries
- Shape and polarity
  - If polar bonds and unsymmetrical, probably polar
- Hybridization
- Valence bond theory: theory based on the idea of electrons in quantum-mechanical orbitals

- Lowest interaction energy (potential energy is at certain distance with substantial overlap but without too much nuclear repulsion

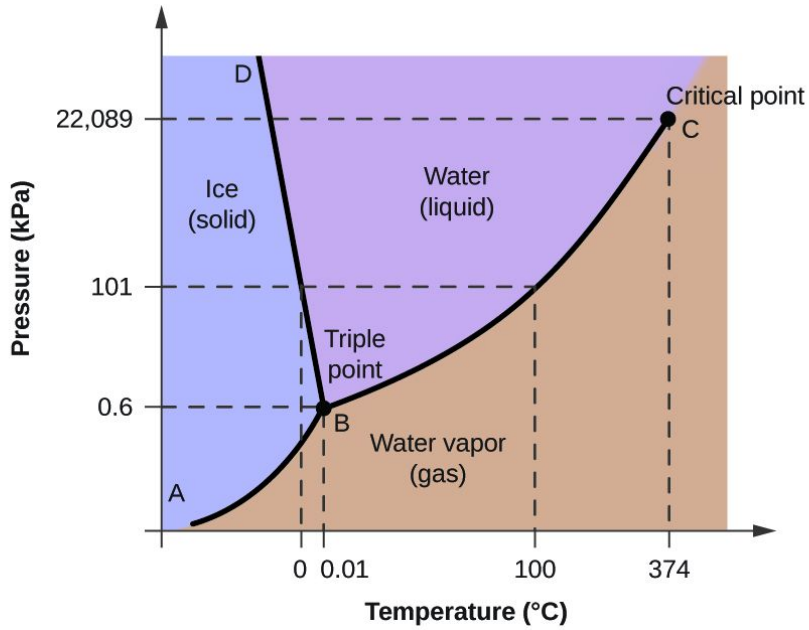


- Two half-filled orbitals overlap to form bond
- Hybridization: degenerate "hybrid" orbitals
  - Minimizes energy by maximizing overlap
  - Depends on number of electron groups (not bonding pairs):
    - Two electron groups: sp
    - 3: sp^2
    - 4: sp^3
    - 5: sp^3d
    - 6: sp^3d^2
  - Rotation of double or triple bonds will break molecular geometry, not allowed
- **Molecular orbital theory: electron delocalization**
- *Lab*: Molecular Modeling

Chapter 11: Liquids, Solids, and Intermolecular Forces

- Solids (crystalline)
- Liquids
- Gases
  - This state can be induced by pressure changes (lower pressure) to solids or liquids, and vice versa
- KMT
  - See Chapter 5
- Phase changes

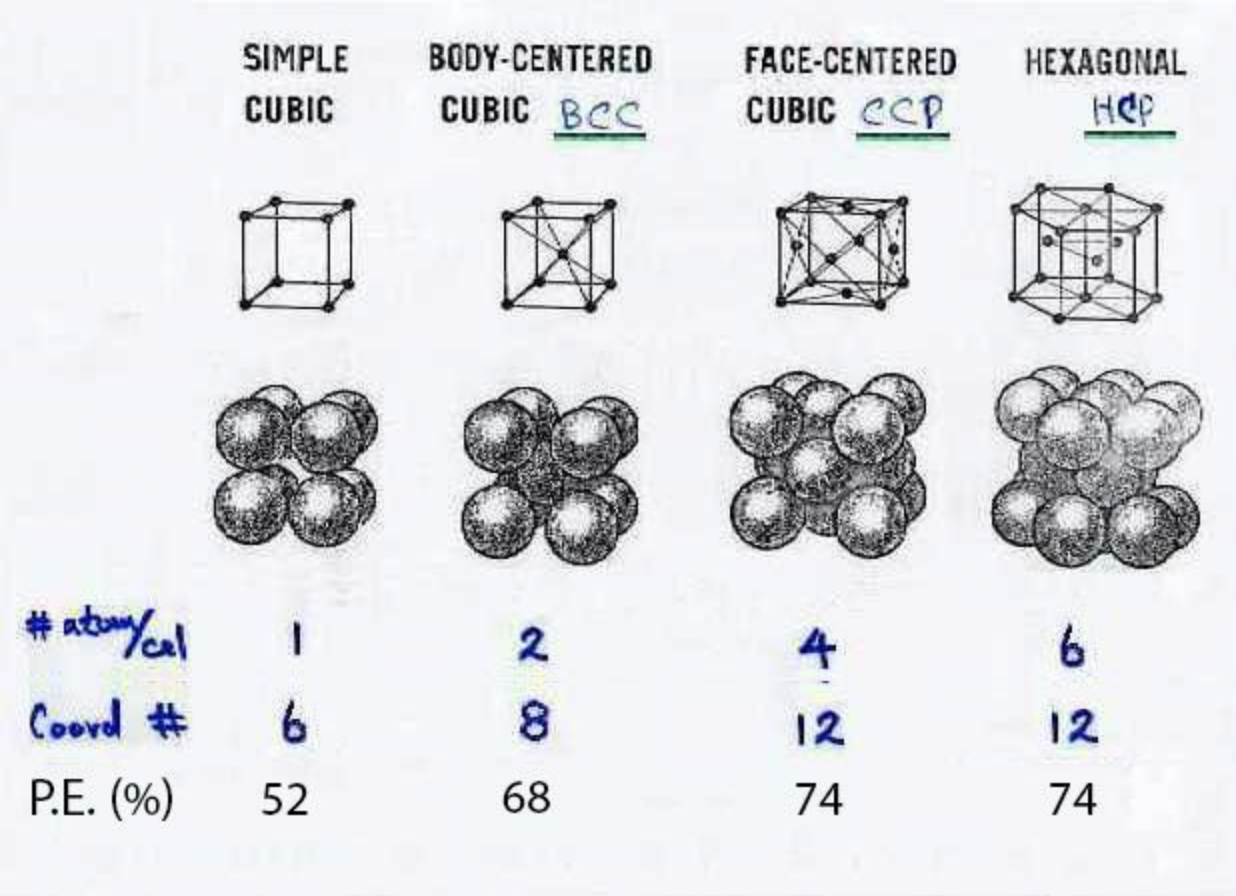
- Vaporization and melting (fusion) are endothermic
- Vaporization continues until dynamic equilibrium is reached
  - Vapor pressure is pressure exerted by vapor of liquid at dynamic equilibrium
    - Vapor pressure correlates exponentially with temperature
      - Clausius-Clapeyron equation: $ln(P_{vap}) = -\frac{\Delta H_{vap}}{R}(\frac{1}{T}) + ln(\beta)$
  - System attempts to keep equilibrium
- Critical point is when supercritical fluid exists
  - Over critical temperature and critical pressure
- Sublimation is direct phase change from solid to gas (and vice versa with deposition)
  - Happen even below boiling point, but often slowly
- Phase diagram represents phase changes at different pressures and temperatures
  - Example:



  - Water's phase diagram is unique because of the positive slope of the liquid/solid line because the liquid is more dense

- Surface tension, capillary action, viscosity
  - Surface tension: the "energy required to increase surface area by a unit amount"
  - Viscosity: ability of liquid to flow
    - Measured in poise (P)
    - Increased by higher intermolecular forces, longer or less regular molecular shape, lower temperatures
  - Capillary action: movement of a liquid up a narrow tube
    - Requires higher adhesive forces than cohesive forces
    - Rises until cohesive and adhesive forces are balanced by gravity
    - Rises higher in a narrower tube
- Intermolecular attractions

    - Because of Coulomb's law, greater distance between molecules than between atoms of molecule make intermolecular attractions much weaker than intramolecular attractions
- X-ray diffraction uses trigonometry and an interference pattern from atoms to determine space between planes
    - Bragg's Law calculates distance between layers using specific wavelength of light and angle
- Regular arrangement of atoms in crystalline solid = crystalline lattice
    - Crystalline lattices can be represented by a unit cell, a repeating structure of a few atoms



    - Packing efficiency is percentage of total space that the atoms occupy
        - Hexagonal closest packing (in hcp and ccp) is the highest packing efficiency
            - Used in hexagonal cubic cell and face-centered (which is like a diagonal hexagonal closest packing) (face-centered is cubic closest packing)
    - There are different types of crystalline solids
        - Molecular solids (e.g., ice)
        - Ionic solids: maximize coordinate # and accommodate charge neutrality
            - Usually the more disproportionate the sizes of the ions, the lower the coordination number
        - Atomic solids:
            - Non-binding atomic solids (only noble gases in the solid phase)
            - Metallic solids

        - Network covalent atomic solids (e.g., diamond or graphite)
- Band theory: a solid bonding model
    - Electrons are delocalized over the entire crystal (not just atom or orbital)
    - Electrons form a band known as the valence band
        - There is also the conduction band formed of the unfilled, higher-energy orbitals that electrons can use to freely move around the crystal
    - The band gap is small in metals, medium in semimetals, and large in nonmetals
        - In semimetals, the band gap can be controlled using doping
            - Elements with more valence electrons can be doped as impurities to create n-type semiconductors (negative charge because more electrons in conductive band), and vice versa for p-type semiconductors
                - Different semiconductors can be combined to form p-n junctions, which are common computer components
                    - P-n junctions can act as diodes or amplifiers in electronics
- *Lab*: Clausius-Clapeyron Equation

Formulas

- Heat of Reaction Shortcuts:
    - Using Hess's Law: $\Delta H^{o}_{rxn} = \sum n_p \Delta H^{o}_{f}(products) - \sum n_r \Delta H^{o}_{f}(reactants)$
    - Using bond energies:

      $$\Delta H^{o}_{rxn} = \sum(bond\ energies\ of\ reactants) - \sum(bond\ energies\ of\ products)$$
- Formal charge: $FC = VE - \frac{1}{2}BE - LE$
    - FC = formal charge
    - VE = valence electrons
    - BE = bonding electrons
    - LE = lone electrons
- Enthalpy: $H = E + PV$
    - H = enthalpy
    - P = pressure
    - V = volume
- Pressure-Volume Work: $w = -P\Delta T$
    - w = work
    - P = pressure
    - T = temperature
- Ideal Gas Law: $PV = n \times RT$
    - P = pressure (atm)
    - V = volume (L)
    - n = moles

    - R = gas constant ((0.08206 L*atm/mol*K)) - depends on which units you use, affects other units in equation for other variables
    - T = temperature (K)
- Van der Waals Equation: $(P + a(\frac{n}{a})^2)(V - nb) = nRT$
    - a, b = constants dependent on substance
    - (rest of constants same as ideal gas law)
    - Know the concept about how real gases compare to ideal (write how here, and why)
- Combined Gas Law: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
    - P = pressure (any unit)
    - V = volume (any unit)
    - T = temperature (K)
- Root Mean Square Velocity: $\mu = \sqrt{\frac{3RT}{M}}$
    - μ = root mean square velocity, used to find velocity
    - R = ideal gas constant
    - T = temperature (K)
    - M = molar mass
- Dipole Moment: $\mu = qd$
    - μ = dipole moment
    - q = charge
    - d = distance
- Probability density: $\psi^2 = \frac{probability}{unit\ volume}$
    - Ψ = probability density
- Radial distribution function: $radial\ probability\ at\ radius\ r = \psi^2 \times (volume\ of\ shell\ at\ r)$
- Coulomb's Law: $F = k \times \frac{q_1q_2}{d^2}$
    - F = electromagnetic attraction between particles
    - k = Coulomb's constant
    - q_1 and q_2 = (signed) charges of particles
    - d = distance between particles
- Change of Molarity Function: $M_1V_1 = M_2V_2$
    - M = molarity
    - V = volume

Constants

- Ideal gas constant (R)
    - $R = \frac{0.08206 atm \times L}{mol \times K}$
    - $R = \frac{8.314 J}{K \times mol}$
- Joule (J)
    - $J = k \times \frac{m^2}{s^2}$
- Avogadro's constant (N_A)
    - $N_A = 6.022 \times 10^{23}$

- Planck's constant
  - 6.626×10^{-34} Js
- Speed of light
  - 3.00×10^{8} m/s

Other things to know

- How to calculate specific heat, or use it
  - Also know other aspects of calorimetry such as c_{cal} (how to calculate?), q_{cal}, etc.
- Basic elements and symbols
- Polyatomic ions

TABLE 2.5 Common Polyatomic Ions

Ion	Name	Ion	Name
Hg_2^{2+}	Mercury(I)	NCS^-	Thiocyanate
NH_4^+	Ammonium	CO_3^{2-}	Carbonate
NO_2^-	Nitrite	HCO_3^-	Hydrogen carbonate (bicarbonate is a widely used common name)
NO_3^-	Nitrate		
SO_3^{2-}	Sulfite		
SO_4^{2-}	Sulfate	ClO^-	Hypochlorite
HSO_4^-	Hydrogen sulfate (bisulfate is a widely used common name)	ClO_2^-	Chlorite
		ClO_3^-	Chlorate
		ClO_4^-	Perchlorate
OH^-	Hydroxide	$C_2H_3O_2^-$	Acetate
CN^-	Cyanide	MnO_4^-	Permanganate
PO_4^{3-}	Phosphate	$Cr_2O_7^{2-}$	Dichromate
HPO_4^{2-}	Hydrogen phosphate	CrO_4^{2-}	Chromate
$H_2PO_4^-$	Dihydrogen phosphate	O_2^{2-}	Peroxide
		$C_2O_4^{2-}$	Oxalate

- Solubility rules

TABLE 3.1 Solubility Guidelines for Common Ionic Compounds in Water

Soluble Ionic Compounds		Important Exceptions
Compounds containing	NO_3^-	None
	$C_2H_3O_2^-$	None
	Cl^-	Compounds of Ag^+, Hg_2^{2+} and Pb^{2+}
	Br^-	Compounds of Ag^+, Hg_2^{2+} and Pb^{2+}
	I^-	Compounds of Ag^+, Hg_2^{2+} and Pb^{2+}
	SO_4^{2-}	Compounds of Sr^{2+}, Ba^{2+}, Hg_2^{2+} and Pb^{2+}
Insoluble Ionic Compounds		**Important Exceptions**
Compounds containing	S^{2-}	Compounds of NH_4^+, the alkali metal cations and Ca^{2+}, Sr^{2+} and Ba^{2+}
	CO_3^{2-}	Compounds of NH_4^+ and the alkali metal cations
	PO_4^{3-}	Compounds of NH_4^+ and the alkali metal cations
	OH^-	Compounds of the alkali metal cations, and NH_4^+, Ca^{2+}, Sr^{2+} and Ba^{2+}

- How to calculate empirical formulas
- How to balance a chemical reaction

- Systematic and hydrocarbon nomenclature