

Chapter 1 Notes

Physical/**Qualitative** Property – Color, smell, hardness, etc.

Measurements/**Quantitative** Measurements – numbers/exact figures

To Infer – to make an observation based on an observation (Beaker contains Water)

Consistent/Artifacts – Consistent numbers will be repeated, Artifacts vary

All non-0 numbers are significant

0's in the **MIDDLE** of a number are **ALWAYS** significant (2048 and .2008 both have 4 sig digits)

0's in **FRONT** of a number are **NEVER** significant

When decimal point is **Present**, 0's **ARE** significant at the END. When **not present**, 0's **ARE NOT**

Put number in Sci Notation to get sig digits

Multiply/Divide – Same # of sig digits as the number in problem with the **LEAST** number of sig digits

Add/Subtract – Last sig digit is the largest uncertain digit.

Conversions – **2.54cm = 1 Inch** **2.20lb = 1kg** **1cal = 4.184 Joules** **1 Cal = 4184J**

% = **Parts over 100**. **PPM = Parts per Million**

Temperature is the measure of the average kinetic energy of the molecules. **0K = -273.15 C Absolute Zero**

Extensive Properties – dependent on amount (mass/volume)

Intensive Properties – Independent on amount (Boiling/melting point, density, etc) **Density =**

Heat = total kinetic energy

Specific Gravity = Density of substance/Density of water

Mass

Volume

Specific Heat – amount of energy needed to raise the temperature of **a gram** of substance by **1 degree C**

Specific Heat of Water = 1 CALORIE

Specific Heat x mass x temperature change = Heat Energy in JOULES. Then Convert to **cal**.

Exothermic/Endothermic – Releases/Absorbs heat.

Physical Properties – can be measured from a sample ALONE. (Density, MP, BP, color, etc)

Chemical Properties – Measured when a sample is mixed with another chemical (reaction in acid, burn in O₂)

Homogeneous/Heterogeneous – Solution, uniform properties throughout/Not uniform distribution (oil/water)

Mixtures – substances that are NOT chemically combined

Filtration – Filtrate is what passes through the filter.

Distillation – Heat a substance to evaporate it (eg, salt/water). Complete separation may not be possible

Chromatography – Differences in solubility VS adhesion to the substrate

Law of Composition/Multiple Proportions – Sample can be broken into %/elements are in whole # ratio

Chapter 2 Notes

Stoichiometry – study of the quantitative relationships in chemical formulas and equations

Gram molecular mass – sum of all atomic masses of atoms written in grams (**units are u or amu for atoms**)

Mole – Chemist's dozen, **6.02 x 10²³** atoms/molecules, 1 mole of any gas has a volume of **22.4 Liters** at STP

Percentage composition – Given % comp, assume 100grams of substance. All % are then changed to grams

Change all weights to MOLES

Setup formula, and divide by the LOWEST value, then multiply to get whole numbers

Gas Density – Molar Mass – Given a gas is 3.165g/Liter at STP, what is the molar mass?

3.165/1 = Mass/22.4 **Setup ratio to convert 1 liter into 22.4 liters (1 mole)**

Common Names – H₂O (Water), NH₃ (Ammonia), CH₄ (Methane)

BE CAREFUL OF MOLES/GRAMS

Diatomic Molecules – Molecules with 2 atoms each. HONBrClIF

Heart of the Problem format - Moles of DESIRED/Moles of GIVEN

Molarity – Moles of Solute/Liters of solution. .5 molarity means .5 moles of solute for 1 liter of solution

Initial Volume x Initial molarity = Final Volume x Final Molarity

Acids form the **H+(Hydrogen)** ion. Bases form **OH- (Hydroxide)** ion. Acids + Bases = Water + Salt

Moles of H+ = Moles in OH- in a neutralization

$$V_a \times M_a = V_b \times M_b$$

Chapter 3 Notes

Protons = +1, 1amu mass **Neutron** = +0, 1 amu mass **Electron** = -1, 1/1837 amu mass

Bohr (Solar-system model) – electrons have certain allowed energy levels **Cavendish** – Neutrons

Atomic # = number of PROTONS **Atomic mass #** = number of protons + neutrons

Isotopes – Same # of protons, different neutrons. Same atomic #, different mass #. Mass#/Atomic# Element

Periodic Atomic Weight is the WEIGHTED AVERAGE of the weights of the isotopes

Democritus (Atomos) – philosopher who decided that matter was discontinuous

Dalton (billiard-ball model) – experiments w/ gases, different substances are different combos of atoms

Tomson (Plum-pudding model) – gas-discharge tubes, atoms have +/- parts, **electron(invent)** equal in all atoms

Rutherford (nuclear model) – most of the mass of the atom is concentrated in positive charged nucleus

Gold Foil Experiment – sent Alpha particles through a sheet of gold, most went through, other bounced

Nucleus positive charged since alpha was repelled, nucleus dense since bounce at **small ref. angle**

Meyer/Mendeleev – Periodic table, predicted undiscovered elements (Ga, Ge, Sc), order by atomic # (TE, I)

Henri Becquerel/Marie Curie – Radioactivity, named by Curie

Alpha – helium, 2 protons/2 neutrons **Beta** – high energy electron (-1) **Gamma** – high energy light

Elements are **metals, nonmetals, and metalloids.**

Metalloids = Boron (B), Silicon (Si), Germanium (Ge), Arsenic (As), Antimony (Sb), Tellurium (Te), Polonium (Po), Astatine (At)

Ductility – the ability to pull a substance into wire

Secility – ability to cut with a knife

Malleability – the ability to pound into sheet (most is gold)

Conductivity – ability to carry electrical current

Horizontal rows are Periods, vertical are groups/families.

Lanthanides are the rare earth metals

IA – Alkali metals **IIA** – Alkali Earth Metals **VIIA** – Halogens **0** – noble gases, **A/0** – Representative

Cathode Ray Tubes – tubes with 2 electrodes, that glow when air is pumped out and voltage is applied

The design of neon lights, cathode (-) rays are **beams of electrons**, same for all substances

Canal rays, the positive ions left after ionizing the gases, differ from substance to substance

The amount of deflection is proportional to its charge (large charge deflects more than small)

The amount of deflection is inversely proportional to mass of particle (large harder to deflect than small)

The final amount of deflection is the e/m (charge – mass) ratio, **Charge of electron -1.60×10^{-19}**

Chapter 4 Notes

Electronic Structure – the way an atom's electrons are arranged **EMR** – X rays, light, IR, UV, Gamma, etc

Amplitude – maximum height of peaks. **Wavelength** – distance between consecutive peaks

Frequency (ν) – number of peaks that pass a given point in 1 second. **Speed of wave = Wavelength * Freq**

Hertz – Hz, number of vibrations/second

Speed of light = 3×10^8

WL = wavelength

Continuous spectrum – all wavelengths of light are present. Each element produces unique set of WL's

Line spectrum/Atomic Emission Spectrum – element that's excited, only certain wavelengths/colors

Rydberg Equation – $1/\lambda = 109,678 \text{ cm}^{-1} * [(1/n_1^2) - (1/n_2^2)]$ n_2 greater than n_1 (if $n_1 = 1$, n_2 equals 2, 3, 4, etc. (Lyman Series) $n_1 = 2$, $n_2 = 3, 4, 5$, etc. (Balmer Series) $n = 3$ (Paschen) $n = 4$ Brackett) Page 109/110

Moseley – dude that concluded that the freq of X rays depended on the material used for the anode

Atomic number represented the # of proton in nucleus. Electrons move w/ acceleration, they give off EMR

Light also has particle properties, sometimes, behaves like tiny packets, called **photons**

Energy of photon = frequency * h (Planck's constant, 6.63×10^{-34} Joule seconds)

Light is emitted by an atom only at certain frequencies, so electron is **quantized**, only has certain energies

Energy of the electron = $-A * (1/n^2)$. $A = 2.18 \times 10^{-18}$ $n = \text{quantum \#, whole \#s only 1 is lowest NRG}$

Wave mechanics – theory explaining the behavior of electrons in atoms

$E = mc^2$

Calculating the wavelength of a PARTICLE. Wavelength = $6.63 \times 10^{-34} / (mv)$

Diffraction – light passes through slit whose width is the same as WL, slit acts as a light source, scatters light

Diffraction pattern – the pattern in which a slit diffracts light

In Phase – two waves add to greater intensity

Out of phase – waves canceled each other out

Node – place of 0 amplitude. **Antinode** – height of amplitude **Standing Waves** – waves w/ stationary nodes

Electron "waves" are standing waves. However, since there are 3 dimensions, 3 quantum numbers

n . – principle quantum number. The greater the n , the greater the average distance of the electron from nucleus.

n starts from 1, 2, 3, to infinity. 1 is the K shell, 2 L, 3M, 4N, etc. Determines number of shells

l – azimuthal quantum number each main shell has 1 or more subshells, starts from 0, 1, 2 – $n-1$ for that shell

Ground State – state of lowest energy.

m – Magnetic Quantum Number – orbital within a particular subshell is distinguished by its value of m . Always in the range of $-l$ to $+l$. If $l = 1$, then 3 orbitals, -1, 0, 1, so 2p orbital. Know electron hotel

Spin = to complete the set of 4 quantum numbers, each orbital can hold 2 electrons, so $+1/2$ or $-1/2$ spin

Pauli Exclusion Principle – no two electrons in any one atom may have all 4 quantum numbers the same

Diamagnetic – no attraction for another magnet **Paramagnetic** – weakly attracted

Ferromagnetic – REALLY attracted. All attractions/non attractions due to electrons spin

Electron configuration – the way the electrons are distributed among the orbital. Read Sec 4.5 – 4.7

As you go down a column, atom size goes up. Left to right, atom size goes down. Read study cards for reason

Lanthanide contraction – Hf is the same size as Zr, even though Hf is lower, due to the decrease in size of the lanthanides. **Positive ions are smaller than normal** (electrons removed to get noble gas config, less electrons, less +/- repulsion, more +/- repulsion, less shielding, smaller atom)

Negative ions are larger (Pick up electrons, more shielding, more +/- repulsion, outer shell increases in size)

Ionization energy – energy required to remove an electron from an isolated gaseous atom in ground state

Ionization energies increase as you get closer to nucleus (more +/- attraction).

Electron Affinity – energy that is released or absorbed when an electron is added

First electron is exothermic due to attraction, 2nd is endothermic due to +/- repulsions. Read pages 138/139

Chapter 5/6

These chapters deal with covalent, molecular, and ionic bonding, as well as molecule shapes, electron dots, etc, that are extremely hard to stick into the computer. Please read chapter 5 and 6 and use the study cards.

Chapter 7 Notes

When in a solvent, the solute is uniformly dissolved.

Arrhenius – first explained electrolytes

Solvent – stuff of largest proportion

solute – everything else.

Water always a solvent

Equilibrium – when neutral – ion and ion – neutral reaction happen at the same speed

Shifted left/right of equilibrium – the location of most of the stuff, dissolved or undissolved.

Filtrate – stuff passing through filter, liquid

precipitate – solid stuff left in filter

AgNO₃ dissociates

Acids and Bases ARE electrolytes

Litmus – blue in base, pink in acid, used as an indicator

Neutralization in aqueous solution – Acid + Base → salt + H₂O

acid salts – result of partial neutralization of polyprotic acid. H in the middle can still react

Arrhenius – need protons, solution

Bronsted – only protons, no solution

Lewis – just electron pairs

Given acid, -1 acid for conj base, given base, +1 for acid.

Amphiprotic/amphoteric – both acid/base(H₂O)

Autoionization reaction – reacts with itself to get ions (H₂O, NH₃, Acetic Acid)

Strong acids are really good proton donors. **Metal Oxide (basic anhydride) + H₂O → Metal hydroxide**

Non-metal oxide (acid anhydride) + H₂O → Acid

Basic/Acid anhydride → Salt

CO₂ is an acid in Lewis because it RECEIVES a pair of electrons.

Reducing agent – itself oxidized, causing other one to be reduced

Oxidizing agent – itself reduced, causing other one to be oxidized

Metals are reducing agents, themselves being oxidized

Elemental nonmetals are oxidizing agents, strongest at top right.

Gain, reduce/oxidize, lose

Metal + HCl → H₂ + metal ionCl₂

Metal + H₂O → Metal Ion + OH⁻ + H₂ 2 metals, exchange signs

Metal + H₂O → metal ion + H₂ + OH⁻

reacts with GROUP I

Non oxidizing acids – the only effective oxidizing agent is H⁺(HCl, H₂SO₄), reacts with **GROUP II**

Metal + H⁺ → metal ion + H₂

Oxidizing Acids – stuff containing H⁺ plus a stronger oxidizing agent (HNO₃, Hot H₂SO₄) **GROUP III**

3Cu + 2 NO₃⁻ + 8 H⁺ → 3Cu²⁺ + 2NO_{2/1} + 4 H₂O

HNO₃⁻ → NO_{2/1} H₂O + metal ion

H₂SO₄ → metalSO₄ + SO₂ + H₂O

Aqua Regia is actually Cl₂

When metals react with O₂, corrodes, few uses. O₂ is an oxidizing agent, combustion is a redox equation

When nonmetals react with O₂, usually burns. CO₂, CO, SO₂, P₄O₁₀ alcohol → CO₂ + H₂O

Combustion – reaction with O₂ that gives off heat/light

NH₃ + H₂O → NH₄OH is a WEAK BASE!!!!!!

2 ionic substances will have double

replacement or metathesis, driven by formation of precipitate, gas, weak electrolyte, or ion

The greater the atom radius, the stronger the acid.

Common indicators – litmus paper, cabbage juice, berry juice, phenolphthalein.

Chapter 8 Notes

In order to have metathesis, must have **driving force of precipitate, weak electrolyte, or gas.**

Precipitates form when ion concentrations are large enough to make the reaction mixture supersaturated

Salt is insoluble if a precipitate will form even in the concentrations of ions are small. **Know solubility rules/gases**

In order to synthesize materials, must form water, solid, or gas. **When desired product is insoluble, use precipitate**

Both reactants must be soluble, but one desired can't be soluble. When desired is solution, pay attention to amount used

Another way is Acid/Base Neutralization. In neutralization, when dealing with quantity, **make the soluble one the limiting reactant.**

To make inorganic salts, react acid and metal carbonate, carbonate turns into gas.

Concentration – to express concentration, use percentage, ppm, molarity, normality,

To convert from mL to grams and vice versa, use the **DENSITY formula (mass/volume)**. Ppm is % * 10⁶

% composition – use 100 grams total (if 5%, then 5 grams solute, 95g solvent).

Molarity – FIND THE BALANCED CHEMICAL FORMULA FIRST. Use coefficients. # of moles * Molarity = conc

.240M solution of Al₂(SO₄)₃. Al concentration = .240 * 2 = .480 M. SO₄ concentration = .240 * 3 = .720 M

Recall how to solve normal **stoichiometry** problems. If you forgot, review chapter 2. If you forgot molarity, read Ch 2

To change from moles to mL or vice versa, use proportion of moles/1000mL or 100mL/moles. **Watch limiting reactants**

Pay attention to concentrations when going from moles to mL. If you got .500M CaCl₂, the Cl⁻ concentration is 1.00M!

When asking ion concentration after reaction finishes, add the two mL values. **Read example 8.15 on page 292**

Chemical analysis – experimental determination of chemical composition. Read 8.16 on Page 293

Remember that moles usually react in equal numbers. If you react 5 moles of SO₄, you'll get back 5 moles of SO₄

For Titration, ALWAYS GET A BALANCED EQUATION. When calculating %, find moles, then grams, then %

of eq = n * number of moles

equivalent weight = mass of 1 mol/n

normality = number of equivalents/1 liter of solution

Normality = n * Molarity

VN = VN

For equivalents, find the ratio in which the two can be converted assuming 1 equivalent weight, then use proportion

In Redox, n = total number of electrons exchanged

Normality gets you equivalents, Molarity gets you moles

Equivalents – Find n , find equivalent weights, proportion it

Molecules such as $\text{Cr}_2\text{O}_7^{2-}$ - exchange 6 e-

When dealing with equivalents and electrons, watch for molecules with 2 or 3 bound together

To prepare solution of given molarity, find the number of equivalents/mole, then proportion with number of total equivalents, then multiply by formula weight to get grams. **WATCH OUT FOR UNITS. MAKE SURE THEY'RE THE RIGHT ONES**

Also remember that electron transfer of 1 or 2 electrons means 1 mol produces 1 or two equivalent moles, can be used.

The sum of all the equivalent weights equal the molecular weight

For all conversions, even if it can be done in head, USE PROPORTION

Chapter 9 Notes

Pressure is inversely proportional with Volume (Boyle's Law), $P_1 \times V_1 = P_2 \times V_2$

1 atm (Standard Atmospheric Pressure) = 760 mmHg/torr = 101.3kPa (Kilo Pascal) = 14.7 psi (lb/inches²)

Add atmospheric pressure to measurements because tire gauges measure ABOVE atmos. Pressure

The temperature in Kelvins is **proportional** to the Kinetic Energy of the gas

Temperature is proportional to Volume, V/T is constant (Charles Law) and Pressure (Gay Lussac's Law)

$PV/T = \text{constant}$, or $P_1 \times V_1/T_1 = P_2 \times V_2/T_2$ STP = Standard temp and pressure = 273K, 1 atm (760 torr)

Ideal Gas Law = $PV = nRT$ $V=L$: $P=\text{atm}$: $T=K$: $n=\#$ of moles, $R=\text{Ideal gas constant } R=0.0821\text{Latm/mol K}$

Dalton Law of Partial Pressure – each gas takes up same amount of space, so total P = sum of all P 's

Air, mixture of mostly O_2 and N_2

Moles O_2 /Moles total = Pressure O_2 /Pressure total

Moles N_2 /Moles total = Pressure N_2 /Pressure total $P_{\text{total}} = P_{\text{O}_2} + P_{\text{N}_2}$

Diffusion/Effusion – Gas in a small area spreads over large area, gas in large area escapes through small area

At the same temp, KE of each gas is the same. $M_a \times V_a^2 = M_b \times V_b^2$ **Density is prop to molecular weight**

Graham's Law – rate of effusion inver. Prop to the Sq Root of density of the gases.

Ideal gases have no volume/attractions, Helium is the most "ideal" gas

$$\frac{\text{Effusion Rate A}}{\text{Effusion Rate B}} = \sqrt{\frac{M_B}{M_A}}$$

Real gas law $(P + N^2a/V^2)(V-nb) = nRT$ a = attraction b = size of the real gas particle

Pressure is prop to the force pushing and inver. Prop. To the area of force. $P=F/A$

Chapter 10 Notes

All matter is composed of small, fast moving particles. Gases, solids, and liquids differ in IMF and compressibility

---**Compressibility** – gasses are compressible, liquids and solids are not

---**Diffusion** – gasses diffuse much quicker than solids/liquids because the new substance can travel relatively farther between collisions than with solids/liquids. Fewer interruptions, greater **Mean Free Path (Average travel time)**

---**Volume** – gases expand, liquid retains volume, solids retain both shape and volume. Solids/liquids have more IMF

Although gases do have IMF's, they have little effect, esp. London and dipole, since the atoms are so spread out

---**Surface Tension** – the amount of energy needed to expand the surface area of a liquid. Molecules on the surface are being pulled downward (3 directions instead of 4) and therefore need energy to rise to the top. Since the molecules at the top have higher PE than those not at the surface, liquids try to lessen surface area, and the shape with the least surface area is a sphere. Surface tension is **indirectly proportional to the temperature** (higher KE, easier to escape), **prop to IMF's**

---**Evaporation** – molecule's escape through the surface as vapor. Need high enough KE to overcome IMF's.

Sublimation – direct from solid – gas. As liquid evaporates, temperature drops (as the molecules with high KE leave, the average KE drops, and temp is prop to KE). **The Evaporation rate is proportional to the temp, surface area, but indirectly prop to the amount of IMF's**

---**Dipole Dipole** – Polar molecules, +/- attract. 1% the strength of normal covalent bonds. In a gas, the molecules are too far apart, so the dipole attract has little or no effect.

---**Hydrogen Bonds** – a really strong Dipole, Hydrogen bond covalently to N, O, and F to form REALLY polar molecules

---**London Forces** – The electron cloud becomes "uneven," or more electrons on one side than on another, called a polarized electron cloud. Creates instantaneous dipoles, but only last for a moment due to the movement of the electrons. Much weaker than other types of bonding. C_8H_{18} has more London force than CH_4 , longer chain, more places to bond. Also, the larger the molecule, the less +/- attraction of electrons, more easily polarized electron cloud, easier to get instantaneous dipoles

---**Heat of Vaporization - H_{VAP}** – to keep the temperature constant when evaporation is taking place, addition heat must be added. H_{VAP} is the amount of energy required to evaporate 1 mol of liquid. $H_{\text{VAP}} = H_{\text{VAPOR}} - H_{\text{LIQUID}}$. Deals with PE, needed to turn liquid – gas. **As IMF's increase, H_{VAP} increases (More attraction, more NRG required)**

---**Vapor Pressure of Liquid – THE VAPOR PRESSURE AT A CONSTANT TEMPERATURE IS INDEPENDENT OF THE VOLUME OR AMOUNT OF LIQUID.** When the volume of the system changes, the rate of evaporation does not, so even though it disturbs the equilibrium, as long as there is water left to reestablish equilibrium, the rate of evaporation will ALWAYS equal the rate of condensation, and the vapor pressure will remain constant. To measure vapor pressure, use a barometer, and add liquid to it, then use the difference in the readings to find vapor pressure. The Vapor pressure is

dependent on the Rate of Evaporation, because if rate of evaporation is low, rate of condensation is also low, so the vapor pressure is low. **It is also indirectly proportional to the IMF's (slow rate of evap, low vapor pressure), and directly proportional to the temperature (higher KE, more molecules escape, more evap, higher pressure).** The Vapor Pressure, if graphed, is a curve that is not infinite. As vapor pressure rises, the gas gets denser, while the liquid becomes less dense. Eventually, they will be equal at the **CRITICAL POINT**, with a CRITICAL PRESSURE. Above this point, everything exists as a gas. To condense, must be under critical point

---La Chatelier's principle – When an equilibrium is disturbed, the system will try to change to regain equilibrium.

---Boiling Point – When bubbles start forming in the water. **Vapor Pressure of water must equal the atmospheric temp.** When something boils, the temperature does not change. Adding more heat just causes more bubbles and causes more evaporation, but temperature doesn't change. **The higher the atmospheric pressure, the higher the boiling point. The higher the IMF, the higher the boiling point (More KE needed to escape and evaporate)**

---Heat of Crystallization – amount of energy needed to freeze 1 mol of liquid. **Heat of Fusion** – amount of energy needed to melt 1 mol of solid. Same number, just one positive, 1 negative. When going from solid – liquid, the molecules move apart slightly, so little energy is needed. To go from liquid to gas, move apart a lot, so more NRG needed

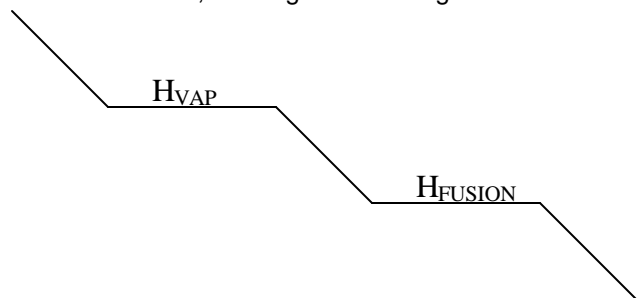
---Molecular crystals – individual atoms at lattice points, London forces, Dipole, H-Bonds. Small lattice energy, easily deformed, soft, low melting points. Bad conductors due to e-'s are bound to molecules and can't move around freely

---Ionic – Ions at lattice points, bonded by electrostatic (non-directional) bonds. When crystals form, they move themselves to maximize attractions and minimize repulsion. Strong electrostatic, high lattice energy, hard, high melting points, brittle, tend to shatter due to ions slip by one another, pattern gets disrupted, starts repelling instead of attracting. Bad conductors in solid form since ions are held in place, but good in aqueous solution

---Covalent – everything bonded to everything else, ultra high melting points, ultra hard, bad conductors

---Metallic – positive ions at lattice points, held by electrostatic between + ions and a sea of electrons. e-' move around freely, so they're good conductors. They range from hard to soft, low to high melting points, and high luster

---When solids reach their melting point, the temperature becomes constant, so no KE change, but PE begins to rise. When all the solid melts, KE begins to rise again.



---Super cooling – for some liquids, the molecules aren't arranged correctly to form crystals. Goes below Freezing Point. When a small crystal gets in place, it acts as a seed, releases the PE, KE rises back to FP, and freezing begins

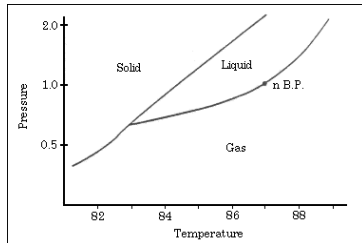
---amorphous solid – a supercooled solid that still moves so slowly they never find their crystal shape. Glass and plastic are two examples

---Vapor Pressure of Solids – similar to those of liquids, but much smaller due to tiny evaporation rate

---Phase Diagrams – points bordering solid – gas, both states, sublimation, two phases at equilibrium

Triple Point – all three phases coexist in equilibrium with each other simultaneously.

On the lines, the two or sometimes three phases are coexisting simultaneously, in equilibrium with each other.



Le Chatelier said that when pressure is increased, it would lead to a denser phase. The rise in pressure leads to the packing of molecules. If one starts at where the solid/liquid are at equilibrium, or at the triple point, and increase the pressure at constant temperature, it should encounter a denser phase. Water, with the curve slanting to the left, is denser as liquid than solid. However, most are like this one, to the right, so solid more dense than liquid.

Questions that I can't answer

Dealing with Vapor Pressure, above the critical point, everything is a gas, right?

Why does the boiling point increase with an increase of IMF's?

Is the critical point of the phase diagram same of the critical point of the vapor pressure. If not, what's the difference?

Chapter 11 Notes

Suspensions – relatively large particles of at least one component are distributed throughout another

Ex, sand in water, blow being blown in the air, precipitate in reaction mixture, settle in influence of gravity

Suspensions don't change **fp/vp**. Molecules are too large, plus too little numbers to have significant effect

Solutions – both solute/solvent are distributed among one another to form single homogeneous phase
 Solution properties different from just solvent properties and solute properties

Colloids — mixtures that are intermediate between solutions/suspensions. Ex, homogenized milk
 Also called **Colloidal dispersions or colloidal suspensions**. Contains mid sized particles
Dispersed phase – stuff being dispersed, “solute” **Dispersing medium** – stuff being dispersed in “solvent”
 Suspended for long periods, *will not separate under influence of gravity?* Can't have gas colloids
Tyndall Effect – scattering of light. **Emulsions** – liquids dispersed in liquids **Stop them from colliding**
Emulsifying Agent – something that forms protective layer around oils/fats so water can mix with them
Adsorption – a process where something sticks to the surface of something else (solid/liquid dispersions)

Substitution solid – solute replaces solvent (ZnS) **Interstitial solid** – Solute squeezes between solvents (WC)

Mole Fraction – moles/total mols **Weight Fraction** – weight/total Weight **Molality** – Moles/1kg

Miscible – dissolvable. When A-B = A-A and B-B, then it's soluble. One polar, one not, not soluble
 For solid to dissolve, must overcome great forces. Molecular stuff is easy to break by non-polar stuff,
 But not polar stuff (polar solvents attract each other too much). If solid is polar or ionic, can't dissolve
 In non-polar solvents (solute sticks to each other too well). Need polar stuff like water(ions)
Solvated – being surrounded by molecules of a solvent. **Hydrated** – surrounded by water
Substances that exhibit SIMILAR IMF's tend to be soluble in one another “Like Dissolves Like”
Hydrophobic – the tail of a “soap” such as Stearate, scared of water. **Hydrophilic** – like water, head
Micelles – collections of fatty acid anions (tails)

Heat of Solution – Amount of energy that is absorbed/released when a substance enters solutions
Exothermic = - number, **Endothermic** = + number Put in Lattice Energy, get back Hydration Energy
Ideal solution – all attractions are equal. Solid same as liquid – lattice first, then hydration
Hydration Energy – net attraction for the solvent dipoles

Solubility – at any temperature, a saturated solution in contact with undissolved solute is at equilibrium (no gas)
Solubility increases with temperature, If you put more solute in it, solubility goes down as temp goes up
 For most solid/liquid substances in a liquid, solubility increases with temperature. **Gases less soluble**
Fractional Crystallization – burn the hell outta it until what you don't want dissolves, then cool it
 Pressure has little to do with liquids/solids. **Solubility of gases increase with increasing pressure**
Henry's Law – Solubility is directly proportional to the PARTIAL PRESSURE. Works for gases only

Vapor Pressure – When making solutions, chemical properties don't change, just physical ones
 *****When a nonvolatile solute is dissolved in liquid, vapor pressure is lowered.*****
Raoult's Law – VP of solution = Mole Fraction of solvent * VP of pure solvent. Nothing to do w/ solute
 When both substances evaporate, get partial pressure of A, then B, add them together
Ideal solutions – follows Rault's Law. **Nonideal** – stuff that doesn't Higher VP, + deviation. Lower, -
 When A-A/B-B less than A-B, - deviation. A-A/B-B more than A-B, + deviation

Colligative Properties – properties that depend on the number of particles instead of their chemical nature
 Ex, Vapor Pressure. Boiling Point elevation – lower VP, higher BP, only for non-volatile substances
 Freezing Point Depression – Freezing point lowered, lower VP, lower FP

Electrolytes – number of moles
 When dealing with VP, when getting mole fraction, especially total moles, take into account NaCl is 2 mols

Rast Problem

$T = K_m$ More concentrated a solution, more off the FP/BP

K- molal FP/BP constant

$M = K * w * 1000 / T * W$

w – weight of unknown solute W – weight of camphor – solvent

T – temperature change K - constant

Chapter 22 Notes

Roentgen – discovered **X Rays, a high energy form of light**

Becquerel – uranium ores emit radiation that passes through objects, **discovered radioactivity**

Marie Curie – Marie/Pierre worked w/ Becquerel. Nobel prize(physics) 1903(shared), and 1911 (radium)

Lawrence – invented cyclotron, used at Berkeley to make transuranium elements

Radioactivity – spontaneous breakdown of atomic nuclei, accompanied by release of some kind of radiation

Half life – the required for half of a radioactive sample to decay. Does not depend on amt of substance

Transmutation – one element being converted into another by a nuclear change

Nuclides – isotopes of elements that are identified by the number of their protons/neutrons

Emmision – the particle ejected from the nucleus

Decay series – the sequence of nuclides that an elements chances into until it forms a stable nucleus

Radioactive dating – using half-life information to determine the age of objects. C-14/C-12 is common

Nuclear fission/fusion – splitting nucleus into smaller parts, smashing two or more nuclei together

Alpha Particles – Helium nucleus, mass of 4, +2 charge, $1/10^{th}$ the speed of light, easy to stop. $4/2He$

Beta Particles – High speed electrons, mass of 0.00055, speed of light, stopped by aluminum, $0/-1e$

Gamma Rays – high energy light, no mass, stopped by several cm of lead
Positron – a positive electron, $0/+1e$ **Neutron** - $1/0n$ $3/0n$ means 3 neutrons
Half Life Problems – Total time, starting amount, half life, ending amount, chart has characteristic curve

Types of decay/why

Too many protons compared to neutrons – positron decay (change proton to neutron and +electron)

Too many neutrons – beta decay (change neutron into a proton and negative Beta particle)

Too many of both – Alpha decay (lose 2 protons and two neutrons)

Radioactive Dating – in every living thing there is a constant ratio of C-12 to C-14, use ratio and half life

Radioisotopes – many substances can be radioactive and then followed as they move through the body

Fission reactors – produce heat which is used to water into steam and drive turbines

Sun/Stars – powered by nuclear fusion – most abundant element in universe is hydrogen, followed by helium

U-235 is fissionable, meaning it can be split when bombarded by neutrons, see **Card 10**

During nuclear change, mass products is less than mass reactants. Called **Mass defect**, used in $E=mc^2$

Moderator – used to slow down neutron flow

Control Rods – Absorbs Neutrons to control reaction speed

\ln ratio of beginning amount/amount final = K (constant) T (time)

B is the middle number(26). **Before is fusion, after is fission**

Chapter 23 Notes

Alkanes = paraffines

alkenes = olefins

aromatic hydrocarbons = stuff with benzene

Ethane = ethylene

propene = propylene

stereoisomerism – geometric/optical

Cis – horizontal **trans** – crisscross

optical isomerism – 4 DIFFERENT atoms to 1 carbon (chiral C atom)

Ethyne – acetylene

alkyl groups – methyl, ethyl, etc

strained – weak bonds in cyclic stuff

Homologous series – each member differs from the next by CH_2 (Alkanes are homologous)

Boat/Chair – 2 isomers of cyclohexane that has no strain

H₂O add 1 H, 1 OH

Ortho/meta/para – 1, 2, or 3 atoms away from the 1st

phenyl – benzene ring as attachment

Markovnikov's rule – hydrogen goes on the carbon with the most hydrogens attacked

Halogen Derivatives – Chloro, Bromo, Fluoro, Iodo

Phenol – alcohol attached to benzene

Methanol = methyl alcohol

carbinol group = C – OH

Cabinol Carbon = the C

Tertiary alcohol – OH attached to a carbon with 3 other carbons attached

Primary Alcohol becomes aldehyde

Keep oxidizing to make keytones and carboxylic acids

Secondary alcohol becomes keytone

Tertiary alcohol doesn't react

Polyhydroxy – 2 or more alcohols

Carbonyl Group = C = O. C+ and O-, turns into ionic group

Aldehydes – “al” ending, C=O on the END

Alcohols – “ol” ending, O – H

Ethanoic Acid = acetic Acid

Acids – “oic” ending, C=O – OH

Keytone – “one” ending, C = O in the middle

Ester – “ate” ending, R – C = O – O – R', combo of an Acid and an Alcohol

Ether – “ether” ending, O in the middle

Amine – “amino + name or Alkyl group + amine” NH_2 , or NH with 2 R's

1 R = primary

2 R's = secondary

3 R's = tertiary

Amide – “name + amide”, C = O – NH_2 Combo of an Acid and an Amine

Peroxide = 2 O's

Monomers – a single piece of a polymer

Addition polymers – using a driving agent, breaking up double bonds (free radicals) and created polymers

Condensation polymers – polymers who join by losing H_2O , used to form esters and amides

Pi bond Shapes – 2 = Linear 3 = boomerang 4 = Y Shaped

Polymerization – to take monomers to form polymers by condensation or addition

Initiation – a double or triple bonded atom is “attacked” by a free radical

Propagation – the free radical breaks the double bond, creates another free radical to attack another molecule

Termination – the free radical encounters another free radical, and they bond, and the chain stops

Free Radical – a lone electron in a half filled orbital

Questions

Why do you need two difunctional monomers for condensation? Isn't it just acid + alcohol?