chemistry final review sheet

Chapter 10: States of Matter

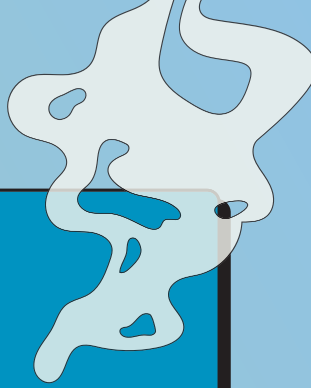
kinetic molecular theory

- used to explain the **properties of solids, liquids, and gases** in terms of the **energy** of particles and the **forces** that act between them
- based on the idea that particles of matter are **always in motion**
- ideal gases are **hypothetical gases** that perfectly fit all assumptions of the kinetic molecular theory

assumptions of kmt

1. particles are **tiny** and **far apart** from each other
   a. most volume occupied by a gas is **empty space**
2. collisions between gas particles and between particles and container walls are **elastic collisions**
   a. no net loss of KE
3. gas particles have **kinetic energy** and are in **continuous, rapid, random motion**
4. **attractive forces** between gas particles **do not exist**
5. gases at same **temperature** have the **same average kinetic energy**

properties



properties of gases

- **expansion: gases do not have definite shape or volume**
    - completely fill an enclosed container
- **fluidity: particles glide past each other because of insignificant attractive forces between particles**
    - both liquids and gases are *fluids*
- **low density: a gas's density equals about 1/1000 of a liquid or solid density of the same substance.**
- **compressibility: decrease space between particles so they move closer together**
- **diffusion: spontaneous mixing of particles caused by their random motion**
    - rapid, random and continuous motion of the gas molecules
- **effusion: gas particles pass through tiny openings**
    - effusion rates of different gases are directly proportional to the velocities of the particles (low mass effuse faster)
- **real gas: gas that does not behave completely according to the assumptions of the KMT**
    - all gases deviate from ideal to some degree
        - more polar = more deviation
    - high temp, low pressure needed to turn gas into ideal gas

properties

properties of liquids

liquids are a state of matter with definite volume that takes the shape of its container

- **attractive forces between particles in a liquid are more effective than those between particles in a gas**
  - attraction between liquid particles are caused by their intermolecular forces (london, dipole-dipole, h-bond)
- **fluid: a substance that can flow and take the shape of its container**
- **high density: 100x denser than gas at normal atmospheric pressure**
- **incompressible: less compressible than gases, more closely packed together**
- **diffusion: liquids gradually diffuse throughout any other liquid in which it can dissolve**
  - much slower than diffusion in gases bc of attractive forces
  - particles are closer together
  - diffusion increases w/ higher temp. due to increased average KE
- **surface tension: force that pulls adjacent parts of a liquid's surface together, decreasing surface area as much as possible**
  - stronger the attractive force, stronger the surface tension
- **capillary action: attraction of the surface of a liquid to the surface of a solid**
  - pulls liquid molecules upward along surface (forms a meniscus)
  - smile meniscus: adhesive forces (sticks to other things) are stronger
  - frown meniscus: cohesive forces (sticks to each other) are stronger

properties

properties of liquids (cont.)

- **vaporization: a liquid or solid changes to a gas**
- **evaporation & boiling: particles escape from the surface of a non boiling liquid and enter the gas state**
  - boiling: change of a liquid to bubbles of vapor that appear throughout the liquid
  - evaporation ONLY occurs at the surface (bc particles of a liquid have different KEs)
- **freezing: the physical change of a liquid to a solid by removal of energy as heat**
  - as liquid is cooled, average energy of particles decreases

properties of solids

- **particles: much more closely packed than liquid or gas**
  - stronger interparticle attraction (they hold particles in relatively fixed positions)
  - more ordered than liquids, much more ordered than gases
- **high density: most substances are densest in the solid state**
- **incompressible: for practical purposes, solids = incompressible**
- **diffusion: the date of diffusion in solids is millions of times slower than in liquids**
- **types: amorphous and crystalline solids**
  - amorphous solids are super cooled liquids and its particles are arranged randomly
  - crystalline solids have orderly, geometric, repeating particles

properties

properties of solids (cont.)

- **melting: the physical change of a solid to a liquid by the addition of energy as heat**
  - definite melting point: temperature at which solid becomes a liquid
  - amorphous solids have no definite melting point (glass, plastic)
    - retain liquid properties even at solid temperatures
- **crystal structure & lattice: the total three-dimensional arrangement of particles of a crystal**

binding forces

crystal structures can be described using the types of binding forces

- **ionic: positive and negative ions arranged in a regular pattern**
  - ionic crystals form when metals combine w/ nonmetals or polyatomic ions
- **covalent network: each atom is covalently bonded to its nearest neighboring atoms**
  - extends throughout a network that includes a very large number of atoms
  - very hard and brittle, high melting points, non- or semi-conductors
- **metallic crystals: metal cations surrounded by a sea of delocalized electrons**
  - freedom of the delocalized electrons explains high conductivity
- **covalent molecular crystals: covalently bonded molecules held together by intermolecular forces**
  - low melting points, easily vaporized, relatively soft, good insulator

changes of state

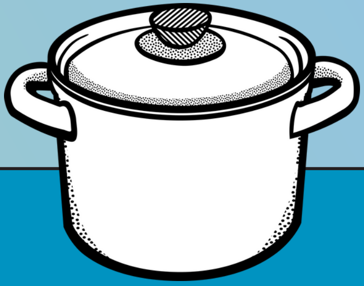
vocabulary

- phase: any part of a system that has uniform composition and properties
- condensation: gas changes to a liquid
- vapor: a gas in contact with its liquid or solid phase
- equilibrium: dynamic condition in which two opposing changes occur at equal rates in a closed system

equilibrium

- when rate of condensation = rate of evaporation, equilibrium is achieved
- static equilibrium: no activity, particles are not in motion
- dynamic equilibrium: moving, particles are in motion
- equilibrium vapor pressure: pressure exerted by a vapor in equilibrium with its corresponding liquid at a given temperature
    - increases with temperature because of increase in KE
    - every liquid has a specific equilibrium vapor pressure at a given temperature
    - equilibrium vapor pressure determines the strength of the binding force (inverse relationship)
- volatile liquids evaporate readily due to relatively weak forces of attraction between their particles
- nonvolatile liquids do not evaporate readily

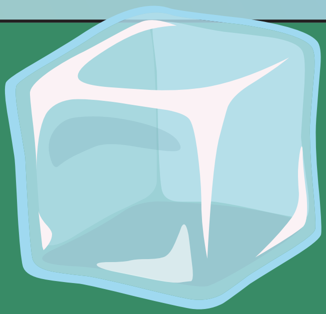
changes of state



boiling

- **conversion of a liquid to a vapor within the liquid as well as at its surface**
- **boiling point: when equilibrium vapor pressure = atm pressure**
  - 1 atm, 760 torr, 101.3 kPa
- **energy must be continually added to keep a liquid boiling**
  - but temperature remains constant bc excess energy breaks bonds
- **molar enthalpy of vaporization: energy released when one mole is vaporized**
  - each liquid has a characteristic enthalpy of vaporization

freezing & melting



- **freezing: physical change from liquid to solid**
- **loss of energy in the form of heat by liquid**
  - liquid = solid + energy
- **change occurs at a constant temperature when it's a pure crystalline substance**
- **freezing point: temp when solid and liquid are at equilibrium (1 atm, 760 torr, 101.3 kPa)**
  - particles of solid and liquid have the same KE
- **melting also occurs at constant temperature**
  - solid + energy = liquid
- **at equilibrium, melting and freezing rates are equal**
- **at normal atmospheric pressure, temperature of a system with ice and liquid will remain at 0°C**

changes of state

sublimation & deposition

- **at sufficiently low temperature and pressure conditions, a liquid cannot exist**
  - a solid substance exists in equilibrium with its vapor instead of liquid
- **change of state from solid to gas is sublimation**
  - ex: dry ice (CO2), iodine
- **reverse of sublimation is deposition (gas to solid)**

phase diagrams

- **graph of pressure versus temperature that shows the conditions under which the phases of a substance exist**
- **triple point: indicates the temperature and pressure conditions at which the solid, liquid, and vapor of the substance can coexist at equilibrium**
- **critical point: indicates the critical temperature and critical pressure**
  - furthest right point on a line segment
- **critical temperature: temperature above which the substance cannot exist in the liquid state**
- **critical pressure: the lowest pressure at which the substance can exist as a liquid at the critical temperature**
- **supercritical fluid: northeast of the critical point, not quite a liquid or a gas**
- **negative slope: less dense as a solid than liquid**
- **positive slope: more dense as a solid than liquid**

Chapter 11: Gas Laws

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important constants/conversions

- Celsius to Kelvin: $K = {}^{\circ}C + 273$ *always convert to Kelvins*
- pressure: 1 atm = 760 torr, 101.3 kPa, 760 mmHg
- STP: molar volume = 22.4 L, pressure = 1 atm, temperature = 0 °C
- $R = 0.0821\ (L*atm)/(mol*K)$ or $R = 62.4\ (L*mmHg)/(K*mol)$

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equations

Dalton's Law of Partial Pressure: $P_{total} = P_1 + P_2 + P_3$

Boyle's Law: $P_1*V_1 = P_2*V_2$ ($k = PV$)

Charles' Law: $V_1/T_1 = V_2/T_2$ ($k = V/T$)

Gay Lussac's Law: $P_1/T_1 = P_2/T_2$ ($k = P/T$)

Combined Gas Law: $(P_1*V_1)/T_1 = (P_2*V_2)/T_2$

Ideal Gas Law: $PV = nRT$; $mm = dRT/p$

Avogadro's Law: $V = kn$

Graham's Law of Effusion/Diffusion: rate of effusion of A/rate of effusion of B = $\sqrt{M_B/M_A}$

Chapter 12: Solutions

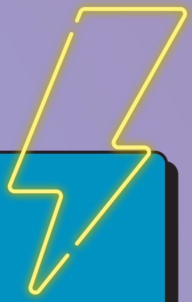
heterogeneous vs. homogenous

- homogenous: fully uniform
- heterogeneous: not fully uniform

solution vs. colloid vs. suspension

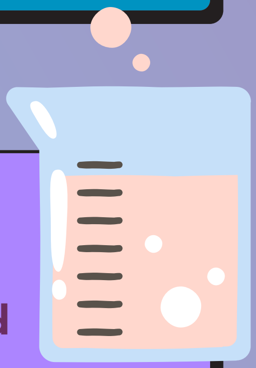
- solution:
    - Does **not separate** on standing or filtration
    - Does **not scatter light**
    - **Clear** (but not colorless)
    - Particles **< 1nm**
- colloids:
    - **heterogeneous**
    - Scatters light (Tyndall effect)
    - Clear or a bit translucent
    - Particles **1-1000nm**
- suspensions:
    - **heterogenous**
    - particles **settle out** and filter out
    - two "**distinct**" phases
    - Particles **> 1000 nm**

solubility



electrolytes

- electrolyte - a substance that dissolves in water and conducts electric current
  - ionic compounds
  - acids
- nonelectrolyte - dissolves in water to give a solution that does not conduct electronic current
  - sugar
- strong electrolytes conduct more electricity than weak electrolytes



rate of dissolution

- to increase rate of dissolution, collisions must be maximized
- increase in surface area of solute
- stirring or shaking of solution
- high temperatures
  - applies only to solids in liquids, liquids in liquids, and liquids in solids
- increase in pressure
  - increases gas solubility
  - increases collisions
  - particles collide with liquid surface causing more gas particles to dissolve in liquid
  - decreasing pressure allows gas to escape solution

solubility

solution equilibrium

- physical state where process of dissolution and crystalization occur at same rates
- solubility depends on:
    - temperature
    - nature of solute/solvent
- unsaturated solution: contains less solute than saturated solution in same conditions
- supersaturated: holding more solute than it should be holding
    - made by increasing temperature then decreasing it slowly
- solubility: g solute per g solvent at specific temperature

"like dissolves like"

- polar dissolves polar
    - ex: acids, water
- nonpolar dissolves nonpolar
    - -ane, -yne, -ene are all nonpolar
- immiscible liquids are not soluble in each other
- miscible liquids dissolve freely in one another

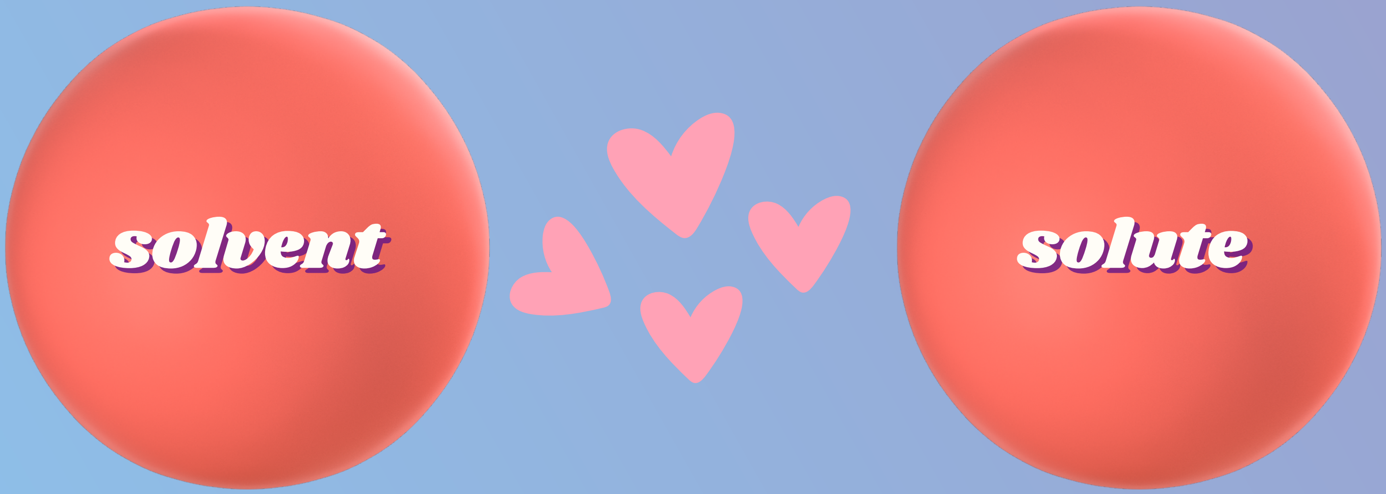
solubility

ionic compound dissolution in H20

- water has a permanent net dipole
- charged parts of water molecules attract ionic compounds and surround them, separating them and drawing them into the solution
- dipole dipole bonding may not be enough to break crystal lattice
    - ex: carbon monoxide

enthalpy of solution

- solute particles are separated from solid (energy is absorbed)
- solvent particles move apart to allow solute to enter solvent (energy absorbed)
- solvent particles are attracted to solute particles (energy released)
    - enthalpy negative when last step is greater than other two
    - exothermic: warms outside, enthalpy negative, more energy released
    - endothermic: cools outside, enthalpy positive, more energy absorbed

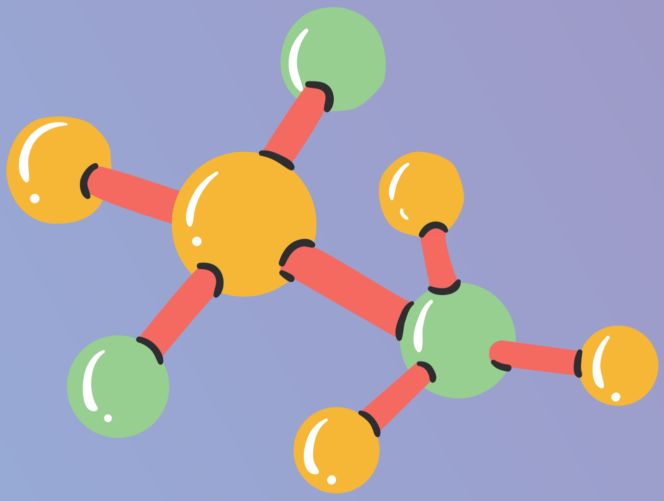


important formulas

concentration

- a measure of solute in a given amount of solvent or solution
- percent by mass: (g solute/g solution)*100
- percent by volume: (mL solute/mL solution)*100
- percent by mass/volume: (g solute/L solution)*100
- molarity (M): (mol solute/L solution)

IMPORTANT



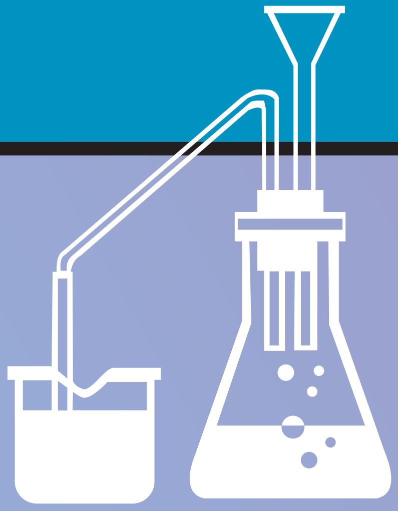
dilutions

- add solvent to solution with too much solute
- concentration vs. dilute
  - concentration --> more solute
  - dilute --> more solvent
- mass (concentrated)*volume (concentrated) = mass(dilute)*volume(dilute)
  - in moles
- molality (m): mol solute/kg solvent
- mole fraction: mol of substance/mol of solution

important formulas

making a solution

- need to know amount and concentration of solution
- use a volumetric flask that's the same volume as volume of solution you need to make
- calculate amount of solute (molar mass of solute * L of solution * M)
- mass amount of solid solute in a balance with a weighing boat
- pour cushion of distilled water into volumetric flask
- use funnel to transfer solid solute to volumetric flask
- use wash bottle to clean off solute stuck to weighing boat, transfer it into the volumetric flask
- cap and swirl/invert volumetric flask
- uncap and fill the volumetric flask with distilled water to the etched line w/ a pipette or dropper
- cap and swirl/invert volumetric flask



spectrophotometry

- spectrophotometer: transmits a beam of colored light to detect amount of light absorbed by sample and amount transmitted
- when things are dissolved in water, their absorption changes and causes different frequencies of lights to be transmitted
- when something absorbs color, it transmits another beam

important formulas

beer-lambert's law

- use to relate absorbance to concentration
- $A = \varepsilon bc$
    - A = absorbance
    - b = path length
    - c = concentration
    - ε = molar absorptivity constant
- linear in most cases except...
    - sample is too concentrated
    - scattering of light
    - if sample fluoresces (produces its own light)
    - if radiation is non-monochromatic
    - if there is stray light
        - from forgetting to close lid

yay! we're done with chapter 12

Chapter 13: Solubility Rules and Colligative Properties

solubility rules

- soluble:
  - group 1 cations (Li+, Na, K, Rb, Cs) and ammonium (NH4)
    - NO exceptions
  - nitrate anion (NO3-), acetate ion (C2H3O2-), chlorate anion (ClO3-), or perchlorate anion (ClO4-)
    - NO exceptions
  - Halides (Cl-, Br, I-)
    - exceptions: silver (Ag+), mercurous/mercuric (Hg), plumbous/plumbic (Pb), calcium (Ca2+), strontium (Sr2+), barium (Ba2+)
- insoluble:
  - carbonate (CO3), phosphate (PO4), chromate (CrO4)
    - exceptions: group 1 cations, ammonium
  - sulfide (S2-)
    - exceptions: group 1 and 2 cations, calcium, strontium, barium, or ammonium
  - hydroxide (OH-)
    - exceptions: group 1 cations, calcium, strontium, barium, or ammonium

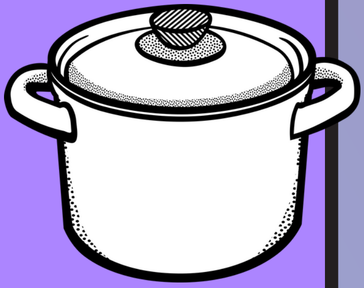
colligative properties

vapor pressure lowering

- vaporization caused by particles leaving from surface
- with more solute, top is being blocked, which lowers amount of stuff trying to leave --> lowers vapor pressure
- Raoult's Law: Psoln = Xsolvent*P0
    - Psoln = pressure of solution
    - Xsolvent = mole fraction of solvent
    - P0 = vapor pressure of the pure solvent

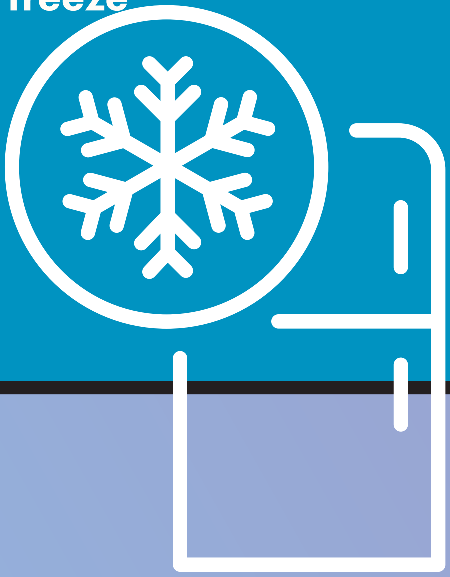
boiling point elevation

- with lower vapor pressure, more energy is needed to boil
- raises boiling point
- change in bp = Kbm
    - Kb = molal boiling point constant
    - m = molality of solution



freezing point depression

- with solute in the way, it becomes difficult to freeze
- need to lose more energy before freezing
- this decreases freezing point
- change in fp = Kfm
    - Kf = molal freezing point constant
    - m = molality of solution



colligative properties

osmotic pressure

- osmosis: water moves from lower solute to higher solute concentration
- osmotic pressure: external pressure needed to stop osmosis
- $\Pi = MRT$
  - Π: osmotic pressure in atm
  - M: molarity of solution
  - R: gas constant
  - T: temperature in Kelvin

Van't Hoff's Factor

- since electrolytes and ionic compounds fully dissociate, they have more than one solute in the solution
- i = number of dissolved particles
  - ex: NaCl --> i = 2
- changes values for colligative properties
  - change in bp = iKbm, change in fp = iKfm, $\Pi = iMRT$

remember how to write net ionic equations!!!

Chapters 14/15: Acids and Bases

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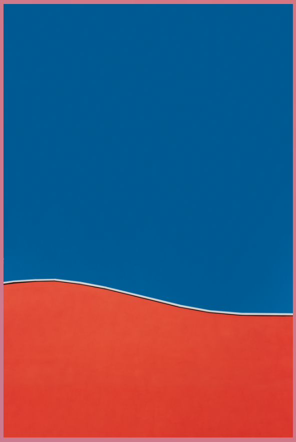
acid vs. base

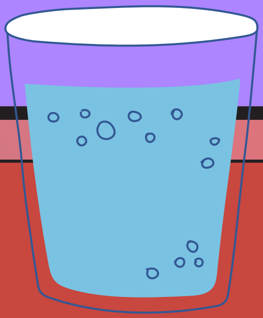
- acids:
    - tase **sour**
    - reacts with **active metals** to form **hydrogen gas + ionic compound**
        - single replacement (use activity series)
        - ex: $6HCl + 2Al \longrightarrow 2AlCl_3 + 3H_2$
    - acids are **electrolytes**
    - change in color of indicator
        - **red** in litmus
        - **colorless** in phenolphthalein
- bases:
    - taste **bitter**
    - feel **slippery** like soap
    - bases are **electrolytes**
    - change in color of indicator
        - **blue** in litmus
        - **pink or magenta** in phenolphthalein
- acids and bases **together form salt**



acids and bases

acid naming conventions

- oxyacid: an acid that contains oxygen
- acid names:
  - ite --> ouse
  - ic --> ate
  - H_2SO_4: sulfuric acid
  - HNO_2: nitrous acid



arrhenius' definition

- acids vs. bases are based on how they react in water
- acid --> increased concentration of $[H^+]$
- base --> increased concentration of $[OH^-]$
- neither means it's neutral

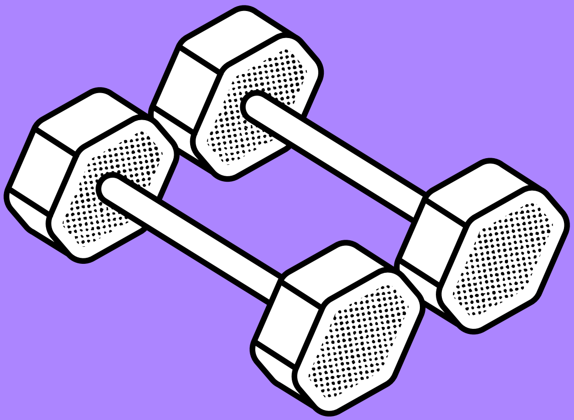
acid/base strength

- strong acid - very low pH (fully dissociates)
  - more conductive
- weak acid - low pH
  - polar
- strong base - very high pH (fully dissociates)
  - more conductive
- weak base - low pH
  - polar

acids and bases

strong acids and bases

- strong acids:
  - Imagine a Perfectly Clear Breezy Summer Night
  - HI - hydroiodic acid
  - HClO4 - perchloric acid
  - HCl - hydrochloric acid
  - HBr - hydrobromic acid
  - H2SO4 - sulfuric acid
  - HNO3 - nitric acid
- strong bases:
  - group 1 and group 2 hydroxides
  - LiOH - lithium hydroxide
  - NaOH - sodium hydroxide
  - KOH - potassium hydroxide
  - RbOH - rubidium hydroxide
  - CsOH - cesium hydroxide
  - Ca(OH)2 - calcium hydroxide
  - Sr(OH)2 - strontium hydroxide
  - Ba(OH)2 - barium hydroxide



bronsted-lowry

- acid: proton donor
- base: proton acceptor
- conjugate base: base that "took" a proton
- conjugate acid: acid that lost a proton
- ex: H2SO4 + H2O --> HSO4- + H3O+
  - H2SO4: acid, H2O: base, HSO4-: conjugate base, H3O+: conjugate acid

acids and bases

lewis

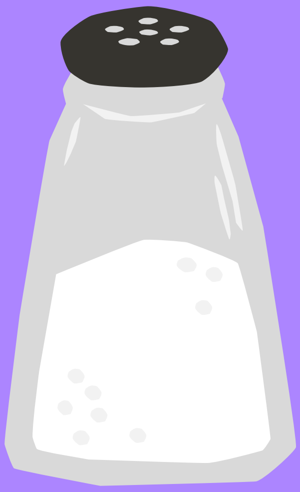
- acids: electron **acceptors**
- bases: electron **donors**
- since there's no protons involved, it applies to reactions that **don't involve hydrogen**
- $BH_3 (g) + NH_3 (g) \rightarrow H_3BNH_3$

amphoteric compounds

- acts as an **acid or base**
  - can **donate or accept** a proton
- has to be **weak**
- can **react with itself**
- $H_2O + H_2O \rightleftarrows OH^- + H_3O^+$

neutralization

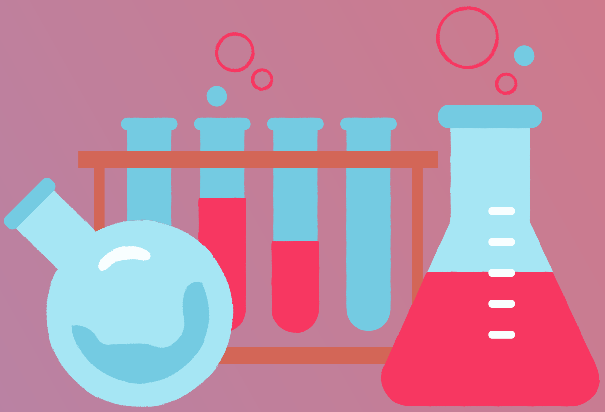
- when acids and base mix, a **salt** is formed
- strong-strong (HCl and NaOH):
  - $H^+ + Cl^- + Na^+ + OH^- \rightarrow Na^+ + Cl^- + H_2O$
  - cancel out: $H^+ + OH^- \rightarrow H_2O$
- weak-strong (HF and NaOH):
  - $HF (aq) + OH^- \rightarrow H_2O (l) + F^- (aq)$
- strong-weak (HCl and NH_3):
  - $H^+ (aq) + NH_3 (aq) \rightarrow NH_4^+ (aq)$
- weak-weak (HF and NH_3):
  - nothing cancels because nothing dissociates



acids and bases

pH and pOH

- pH > 7: basic
- pH < 7: acidic
- pH = 7: neutral
- pH = -log[H+]
- pOH = -log[OH-]
- pH + pOH = 14
- [H+][OH-] = 1.0 * 10^-14
- sigfig rule:
    - sigfigs in concentration is **equal to number of decimal points in pH/pOH**
    - [H+] = 1.0 * 10^-3: 2 sigfigs
    - pH = 3.00: 2 decimal places



indicators

- indicators are **weak** acids/bases
    - **different color** than conjugate base
- transition interval: each indicator has a pH range of values in which it will **change color**

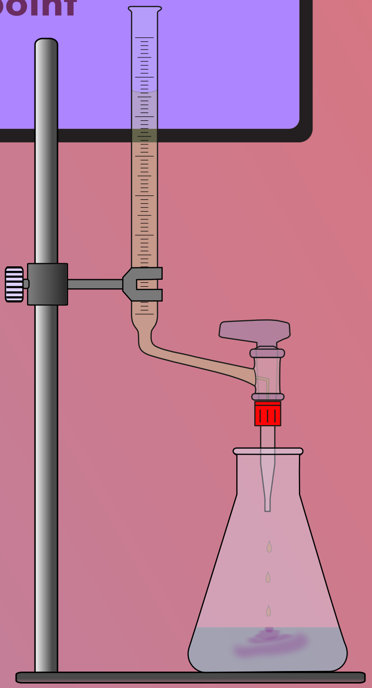
acids and bases

titration

- controlled addition of an acid/base of known concentration to a known volume of acid/base with unknown concentration

STEPS:

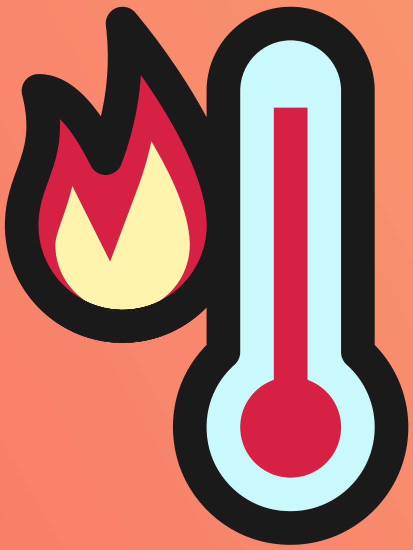
1. take disposable pipette and add about an inch of base into buret, tilt slightly
2. swirl buret to coat entire buret w/base
3. repeat three times (inch of base, swirl, dump into sink)
4. repeat process with acid and volumetric pipet (three times)
5. take erlenmeyer flask, add water cushion
6. use volumetric pipet to put 10 mL of acid of unknown concentration in flask
7. add two drops of indicator
8. swirl to ensure complete dispersion
9. introduce base in buret
10. open valve and get rid of air bubbles
11. close buret w/stopper
12. record initial volume of base
13. record final volume once a color change appears (endpoint happened)



Chapter 16: Reaction Energy

important terms/units

- temperature: measure of average kinetic energy
  - in Kelvin or Celsius
- energy: capacity to do work or produce heat
- heat (q): energy transferred to motion of atoms and molecules
  - transfer of energy from high to low temps
  - Joule (J) = kg*m2/s2
  - calorie (c): amount needed to raise temperature of 1 gram of water by one degree Celsius
  - Calorie (C): 1000 c = 1kcal
  - 1 calorie = 4.184 J
  - 1 Calorie = 4184 J
- work: energy transferred to the motion of objects



reaction energy

enthalpy change

- amount of energy absorbed by a system at constant presure --> measure via change in temperature
- physical and chemical:
  - enthalpy of vaporization or fusion: physical change
  - enthalpy of reaction: chemical change
- state function: depends on only beginning and end
  - ONLY initial and final states are important, path in between is irrelevant

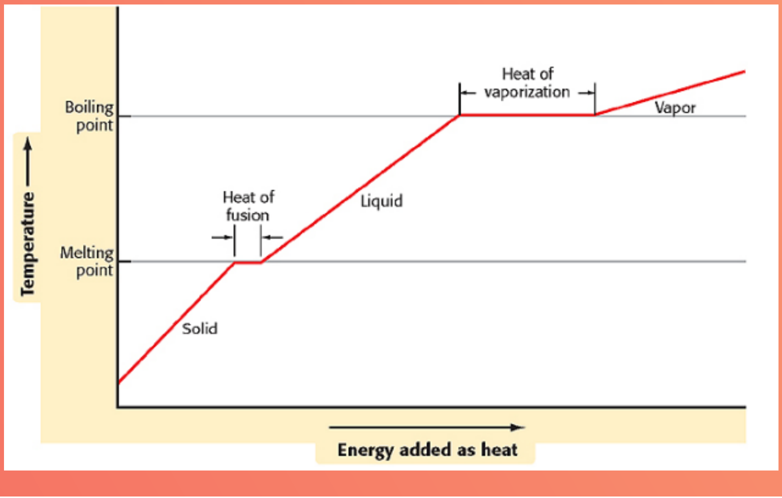
specific heat capacity

- amount of heat it takes to change the temperature of one gram of a substance by 1 degree Celsius
- $q = mc\Delta T$
  - q = heat
  - m = mass in grams
  - c = specific heat capacity
  - ΔT = change in temp (final - initial)
- calorimetry: process of measuring amount of heat released or absorbed during a chemical reaction
  - $-(q_{lost}) = q_{gained}$
  - $-(mc\Delta T) = mc\Delta T$

reaction energy

change of state

- exothermic:
    - heat flows out
    - bonds are formed
    - freezing, condensation, deposition
- endothermic:
    - heat flows in
    - bonds are broken
    - melting, vaporization, sublimation
- heat of fusion:
    - melting/freezing
        - energy needed to go from solid to liquid and vice versa
    - J/g or J/mol
    - ΔHfus
    - first plateau in graph
- heat of vaporization:
    - vaporization/condensation
        - energy needed to go from liquid to gas and vice versa
    - J/g or J/mol
    - ΔHvap
    - second plateau in graph



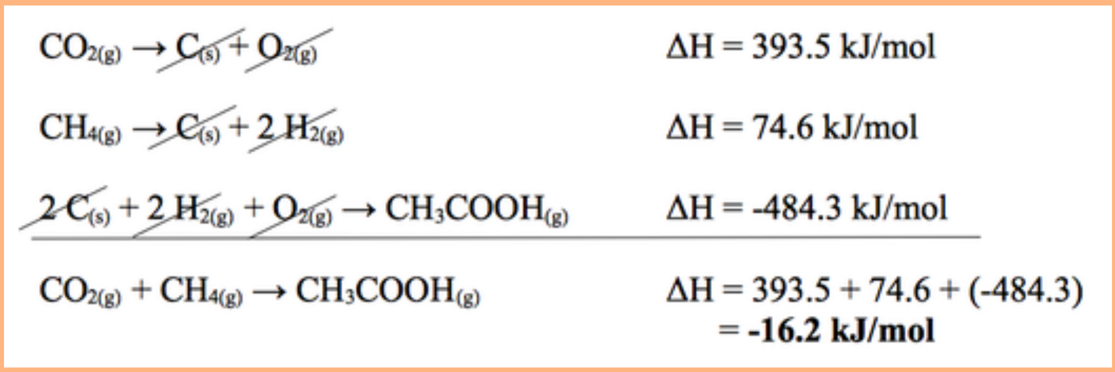
reaction energy

thermochemical equations

- find change in enthalpy
    - pressure is constant
- ΔH = q
    - ΔH - change in heat of reaction
- enthalpy: heat gained or lost by the system
- if ΔH is negative, heat energy is released
- standard state:
    - 25 degrees C, 298 K, 1 atm, 101.3 kPa
- enthalpy change for reaction (ΔH): sum of products - sum of reactants
- standard enthalpy of formation (ΔHf) - energy change for formation of one mole of a compound from its elements at standard state
- enthalppy of combustion (ΔHc) - energy change when one mole of a compound reacts with O2
- the more stable it is, the more energy it takes to break

hess' law

- crossing out/switching of chemical formulas to find ΔH that we want



$$\begin{array}{ll}
CO_{2(g)} \rightarrow \cancel{C_{(s)}} + \cancel{O_{2(g)}} & \Delta H = 393.5 \text{ kJ/mol} \\
CH_{4(g)} \rightarrow \cancel{C_{(s)}} + \cancel{2\,H_{2(g)}} & \Delta H = 74.6 \text{ kJ/mol} \\
\cancel{2\,C_{(s)}} + \cancel{2\,H_{2(g)}} + \cancel{O_{2(g)}} \rightarrow CH_3COOH_{(g)} & \Delta H = -484.3 \text{ kJ/mol} \\
\hline
CO_{2(g)} + CH_{4(g)} \rightarrow CH_3COOH_{(g)} & \Delta H = 393.5 + 74.6 + (-484.3) \\
 & \quad = \mathbf{-16.2\ kJ/mol}
\end{array}$$

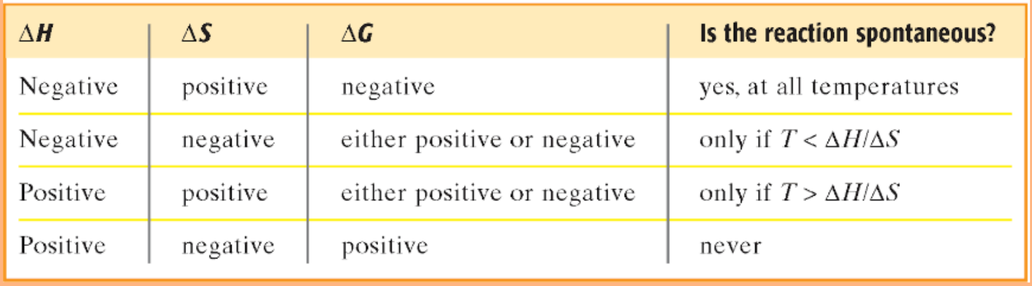
reaction energy

entropy

- measure of relative disorder/chaos (S)
- Sgas >>> Sliquid > Ssolid, Saq > Ss
- state function
  - $\Delta S°$reaction = ΣnpS°products - ΣnrS°reactants
  - $\Delta S > 0$: increasing disorder
  - $\Delta S < 1$: decreasing disorder
- low pressure leads to higher entropy
- more bonds --> less chaos --> less entropy

gibbs free energy

- amount of energy in a system **available to do useful work** (g)
- ΔG = ΔH (kJ/mol) - T (K) * ΔS (J/mol*K)
  - note: parentheses denote units, not multiplication
  - if **$\Delta H > T\Delta S$**...enthalpy-driven
    - $\Delta G > 0$, process **not spontaneous**
  - if **$\Delta H < T\Delta S$**...entropy-driven
    - $\Delta G < 0$, process is **spontaneous**
- -ΔG --> free energy **decreases**, **spontaneous** reaction
- +ΔG --> free energy **decreases**, **spontaneous** reaction
- 0 ΔG: **equilibrium state,** no discernable change in either direction



ΔH	ΔS	ΔG	Is the reaction spontaneous?
Negative	positive	negative	yes, at all temperatures
Negative	negative	either positive or negative	only if $T < \Delta H/\Delta S$
Positive	positive	either positive or negative	only if $T > \Delta H/\Delta S$
Positive	negative	positive	never

Chapter 17: Reaction Kinetics

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important terms/units

- reaction mechanism:
    - **series of steps** in complex reaction
    - **same beginning and end** substances
    - determined **experimentally**
- rate-determining step:
    - **slowest** part of reaction
    - sets **pace** for reaction
- intermediate:
    - substance produced in a mechanism
    - appears as a product, later as a reactant
    - **does not appear** as reactant of final products
- catalyst: **speeds up** reaction
    - **increases** rate of reaction by **targeting rate-determining step**
    - does not appear in overall final reaction

$\text{Cl} + \text{O}_3 \text{ --> } \text{O}_2 + \text{ClO}$

$\text{O}_3 \text{ --> } \text{O}_2 + \text{O}$

$\text{Cl} + \text{O} \text{ --> } \text{Cl} + \text{O}_2$

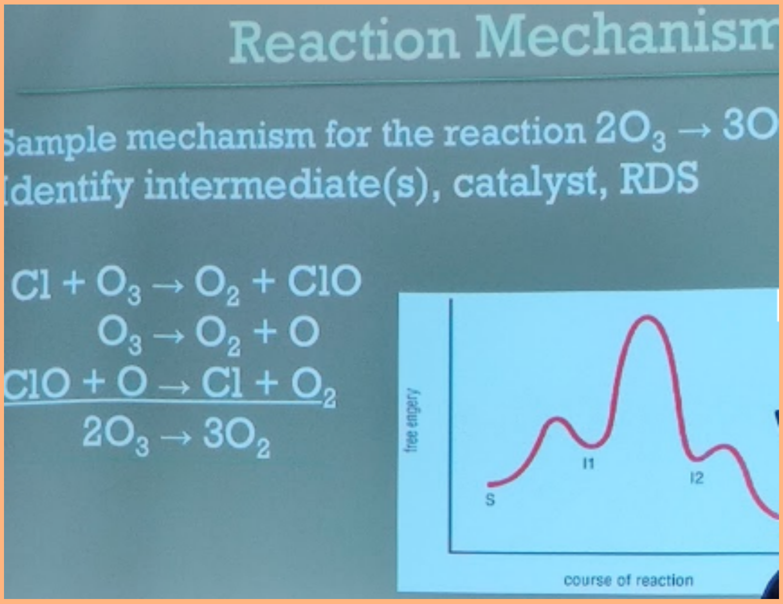
$2\text{O}_3 \text{ --> } 3\text{O}_2$

reactants: O_3, product: O_2, catalyst: Cl, intermediate: ClO

reaction kinetics

reaction mechanism graph

- exothermic: less energy in products than reactants
- # steps: 3 --> number of peaks
- rate determining step: second peak because it has largest energy difference
- activation energy: energy difference from previous valley to top of hill



Reaction Mechanism

Sample mechanism for the reaction $2O_3 \rightarrow 3O$
Identify intermediate(s), catalyst, RDS

$Cl + O_3 \rightarrow O_2 + ClO$
$O_3 \rightarrow O_2 + O$
$ClO + O \rightarrow Cl + O_2$
$2O_3 \rightarrow 3O_2$

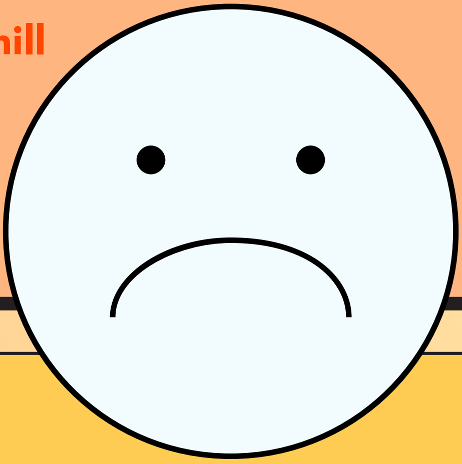
collision theory

- molecules collide for reactions to occur
- two requirements:
  - must have sufficient energy or else they bounce off
  - must approach in the correct orientation
  - collisions that are too gentle or in poor orientation will not result in a reaction

reaction kinetics

activated complex (transition state)

- short-lived substances (old bonds broken and new bonds formed)
- transition state lives less than intermediate state
- so fragile it could fall apart
- could easily go backwards and form reactants
- activation energy (Ea): amount of energy needed for an effective collision to lead to the transition state
  - used to collide particles and achieve a transition state
  - measured from bottom of valley to top of hill
  - slow step has greatest Ea magnitude
  - graph goes down: exothermic
  - graph goes up: endothermic



reaction rate

- chemical kinetics: branch of chemistry that studies reaction rate and mechanisms

these factors affect reaction rate...

- nature of reactants
  - group 1 metals are more reactive than group 2 metals so activation energy is lower for collision
- concentration:
  - more concentration = more collisions = more effective collisions
- surface area:
  - larger surface area = more collisions
- temperature:
  - increase temperature = move faster = more collisions
- catalysts decrease activation energy, which lowers the "hill"

reaction kinetics

catalysts

- **reduce activation energy** for a chemical reaction
- homogenous: catalysts are in **same** state as reactants
- heterogeneous: catalysts are in a **different** state than reactants
    - if there are multiple reactants in different states, we **cannot categorize them** into homogenous or heterogeneous

rate law

- rate law = k[A]^m[B]^n
    - k = rate constant for reaction (determined experimentally)
    - A and B are reactants
    - reaction order: exponents (m and n)
    - total order: sum of reaction order
        - 1st order: when exponent = 1 (reaction occurs in FIRST order w/ respect to exponent's base)
            - direct relationship between concentration and rate law
        - 2nd order: when exponent = 2 (reaction occurs in SECOND order w/ respect to exponent's base)
            - square relationship w/ concentration
        - 0th order: when exponent = 0 (reaction occurs in the ZEROTH order w/ respect to exponent's base)
            - does not relate to the experiment
    - found by dividing known rates

reaction kinetics

rate constant (k)

- once orders are known, **plug in** to find k
- units **depend on order**
  - keep rate in **M/s**
- temperature **dependent**
  - temp increase --> k increases and vice versa

validity

- elementary steps **must sum to overall reaction**
  - intermediates and catalysts **cancel out**
  - coefficients must be correct
- rate law of slow step (rate-determining step) **must equal overall rate law**