

***chemistry final
review sheet***

Chapter 10: States of Matter

kinetic molecular theory

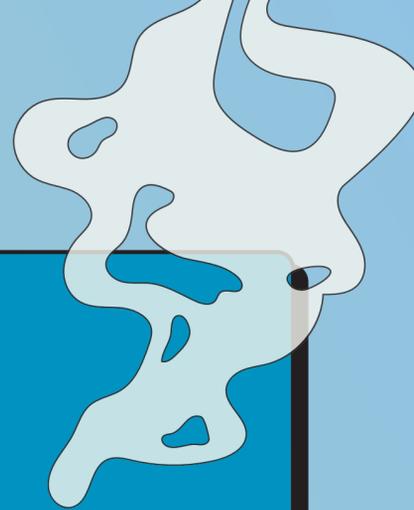
- used to explain the **properties of solids, liquids, and gases** in terms of the **energy** of particles and the **forces** that act between them
- based on the idea that particles of matter are **always in motion**
- ideal gases are **hypothetical gases** that perfectly fit all assumptions of the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



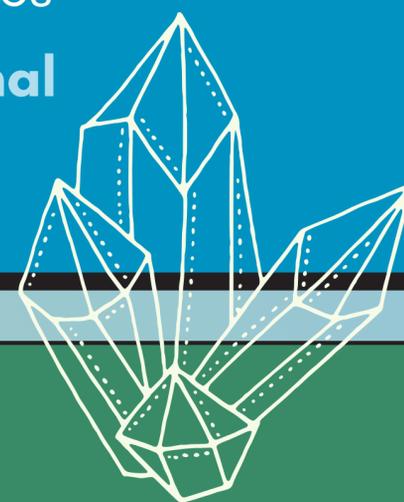
- **particles: much more closely packed than liquid or gas**
 - stronger interparticle attraction (they hold particles in relatively fixed positions)
 - more ordered than liquids, much more ordered than gases
- **high density: most substances are densest in the solid state**
- **incompressible: for practical purposes, solids = incompressible**
- **diffusion: the rate 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**

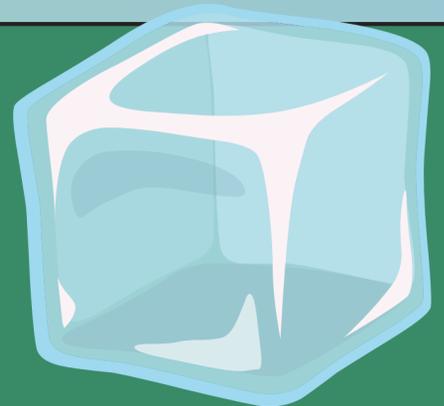
changes of state



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 (CO_2), 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

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
- $R = 0.0821 \text{ (L} \cdot \text{atm)} / (\text{mol} \cdot \text{K})$ or $R = 62.4 \text{ (L} \cdot \text{mmHg)} / (\text{K} \cdot \text{mol})$

equations

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

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

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: $(P_1 \cdot V_1)/T_1 = (P_2 \cdot V_2)/T_2$

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

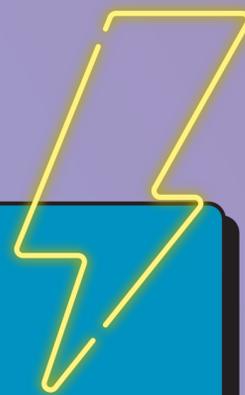
heterogeneous vs. homogenous

- 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 crystallization 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

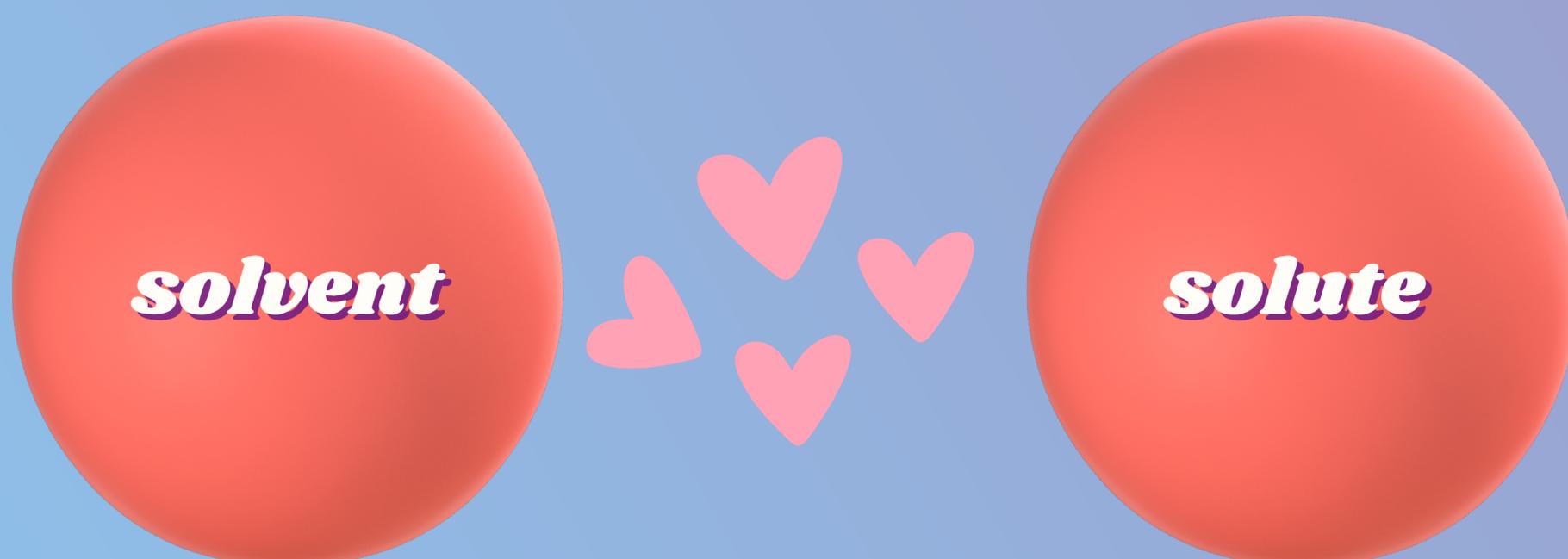
solubility

ionic compound dissolution in H₂O

- 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

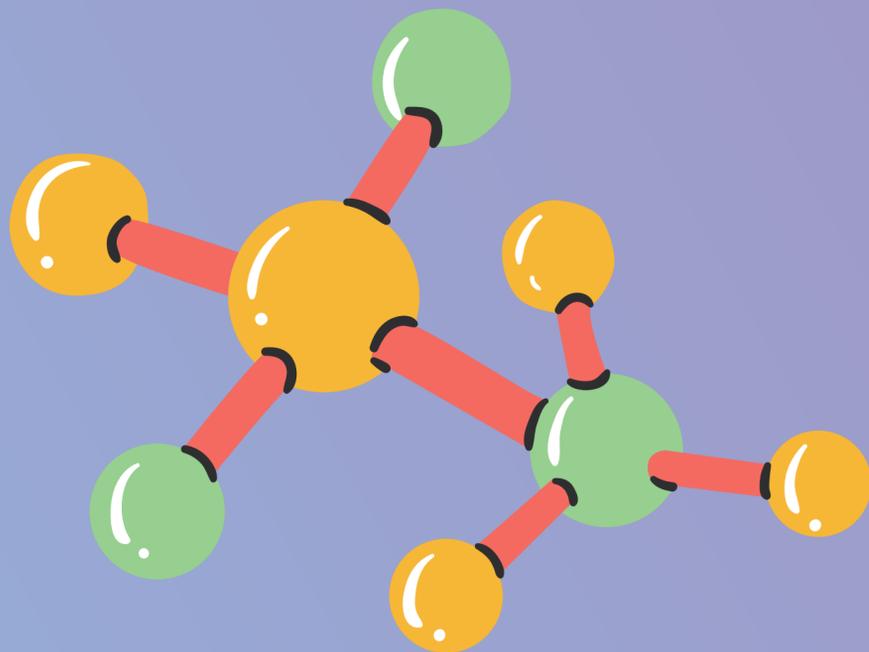


important formulas

concentration

- a measure of solute in a given amount of solvent or solution
- percent by mass: $(\text{g solute}/\text{g solution}) \times 100$
- percent by volume: $(\text{mL solute}/\text{mL solution}) \times 100$
- percent by mass/volume: $(\text{g solute}/\text{L solution}) \times 100$
- molarity (M): $(\text{mol solute}/\text{L solution})$

IMPORTANT



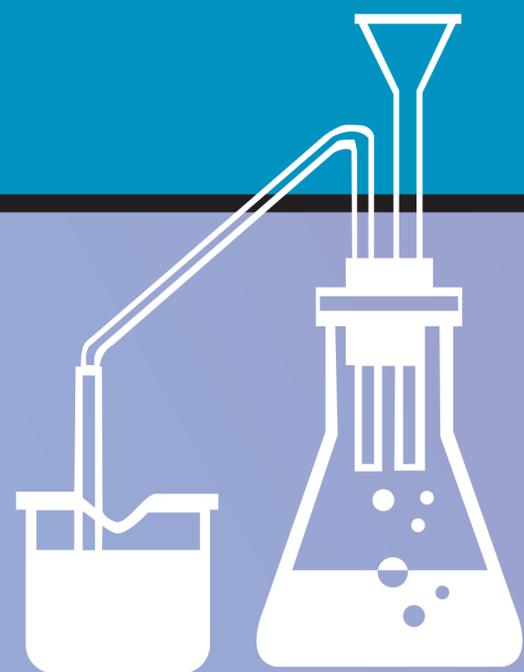
dilutions

- 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
- $A = \epsilon bc$
 - 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** (NH_4)
 - **NO exceptions**
- **nitrate anion** (NO_3^-), **acetate ion** ($\text{C}_2\text{H}_3\text{O}_2^-$), **chlorate anion** (ClO_3^-), or **perchlorate anion** (ClO_4^-)
 - **NO exceptions**
- **Halides** (Cl^- , Br , I^-)
 - **exceptions: silver** (Ag^+), **mercurous/mercuric** (Hg), **plumbous/plumbic** (Pb), **calcium** (Ca^{2+}), **strontium** (Sr^{2+}), **barium** (Ba^{2+})

- **insoluble:**

- **carbonate** (CO_3), **phosphate** (PO_4), **chromate** (CrO_4)
 - **exceptions: group 1 cations, ammonium**
- **sulfide** (S^{2-})
 - **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: $P_{\text{soln}} = X_{\text{solvent}} \cdot P_0$
 - P_{soln} = pressure of solution
 - X_{solvent} = mole fraction of solvent
 - P_0 = 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 = $K_b m$
 - K_b = 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 = $K_f m$
 - K_f = 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: $\text{NaCl} \rightarrow i = 2$
- changes values for colligative properties
 - change in $b_p = iK_b m$, change in $f_p = iK_f m$, $\Pi = iMRT$

remember how to write net ionic equations!!!

Chapters 14/15: Acids and Bases

acid vs. base

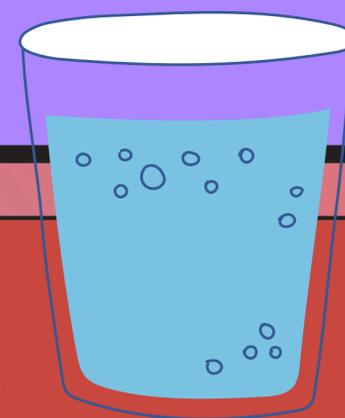
- acids:
 - taste **sour**
 - reacts with **active metals** to form **hydrogen gas + ionic compound**
 - single replacement (use activity series)
 - ex: $6\text{HCl} + 2\text{Al} \rightarrow 2\text{AlCl}_3 + 3\text{H}_2$
 - acids are **electrolytes**
 - change in color of indicator
 - **red** in litmus
 - **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
 - H₂SO₄: sulfuric acid
 - HNO₂: 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

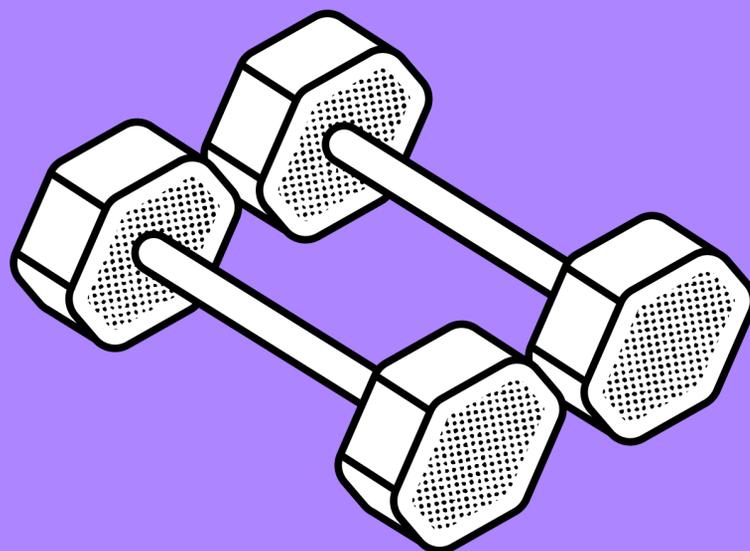
strong acids and bases

- 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

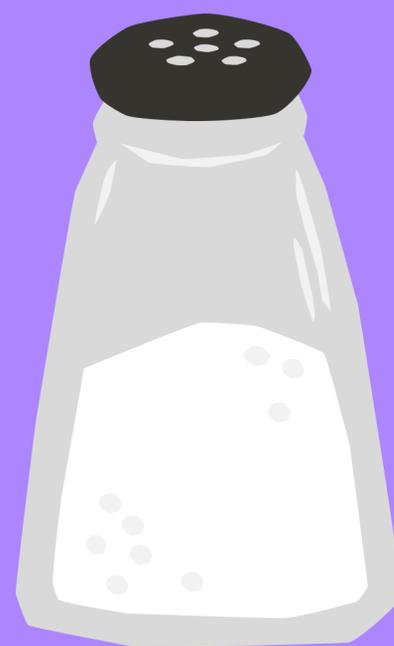
- acids: electron acceptors
- bases: electron donors
- since there's no protons involved, it applies to reactions that don't involve hydrogen
- $\text{BH}_3 (\text{g}) + \text{NH}_3 (\text{g}) \rightarrow \text{H}_3\text{BNH}_3$

amphoteric compounds

- acts as an acid or base
 - can donate or accept a proton
- has to be weak
- can react with itself
- $\text{H}_2\text{O} + \text{H}_2\text{O} \rightleftharpoons \text{OH}^- + \text{H}_3\text{O}^+$

neutralization

- when acids and base mix, a salt is formed
- strong-strong (HCl and NaOH):
 - $\text{H}^+ + \text{Cl}^- + \text{Na}^+ + \text{OH}^- \rightarrow \text{Na}^+ + \text{Cl}^- + \text{H}_2\text{O}$
 - cancel out: $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
- weak-strong (HF and NaOH):
 - $\text{HF} (\text{aq}) + \text{OH}^- \rightarrow \text{H}_2\text{O} (\text{l}) + \text{F}^- (\text{aq})$
- strong-weak (HCl and NH_3):
 - $\text{H}^+ (\text{aq}) + \text{NH}_3 (\text{aq}) \rightarrow \text{NH}_4^+ (\text{aq})$
- weak-weak (HF and NH_3):
 - nothing cancels because nothing dissociates



acids and bases

pH and pOH

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

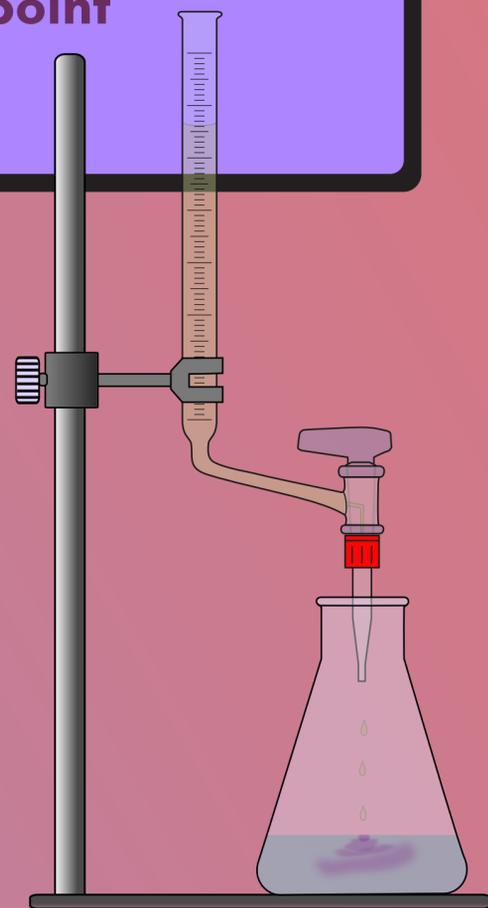
acids and bases

titration

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

STEPS:

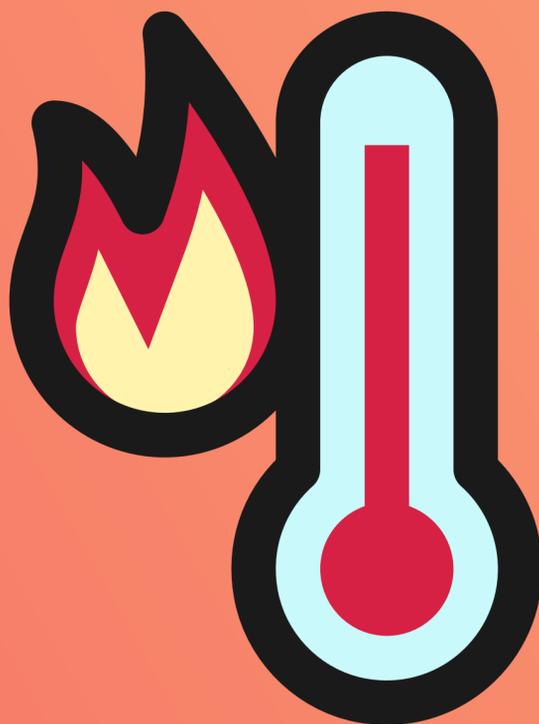
1. **take disposable pipette and add about an inch of base into buret, tilt slightly**
2. **swirl buret to coat entire buret w/base**
3. **repeat three times (inch of base, swirl, dump into sink)**
4. **repeat process with acid and volumetric pipet (three times)**
5. **take erlenmeyer flask, add water 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) = $\text{kg} \cdot \text{m}^2/\text{s}^2$
 - **calorie (c): amount needed to raise temperature of 1 gram of water by one degree Celsius**
 - Calorie (C): $1000 \text{ c} = 1\text{kcal}$
 - **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 pressure --> 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

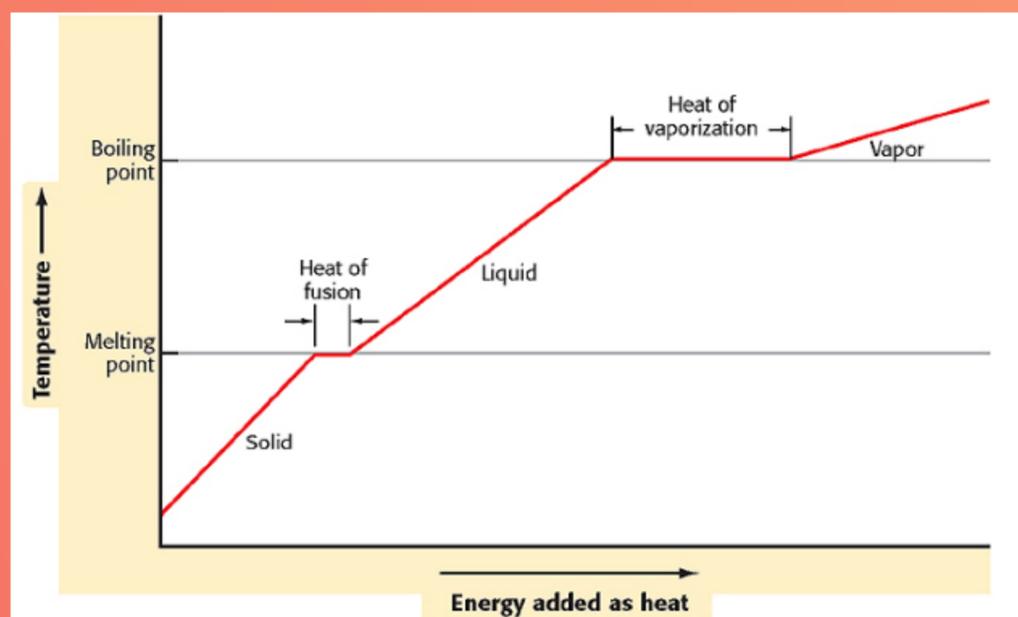
- amount of heat it takes to **change the temperature** of one gram of a substance by **1 degree Celsius**
- **$q = mc\Delta T$**
 - q = heat
 - m = mass in grams
 - c = specific heat capacity
 - ΔT = change in temp (final - initial)
- calorimetry: process of **measuring amount of heat released or absorbed**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



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
 - ΔH_{fus}
 - 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
 - ΔH_{vap}
 - second plateau in graph



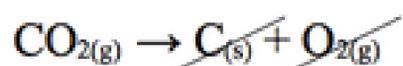
reaction energy

thermochemical equations

- find change in enthalpy
 - pressure is constant
- $\Delta H = q$
 - 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 (ΔH_f) - energy change for formation of one mole of a compound from its elements at standard state
- enthalpy of combustion (ΔH_c) - 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



$$\Delta H = 393.5 \text{ kJ/mol}$$



$$\Delta H = 74.6 \text{ kJ/mol}$$



$$\Delta H = -484.3 \text{ kJ/mol}$$



$$\Delta H = 393.5 + 74.6 + (-484.3) \\ = -16.2 \text{ kJ/mol}$$

reaction energy

entropy

- measure of relative disorder/chaos (S)
- $S_{\text{gas}} \gg S_{\text{liquid}} > S_{\text{solid}}, S_{\text{aq}} > S_{\text{s}}$
- state function
 - $\Delta S^{\circ}_{\text{reaction}} = \sum n_{\text{p}} S^{\circ}_{\text{products}} - \sum n_{\text{r}} S^{\circ}_{\text{reactants}}$
 - $\Delta S > 0$: increasing disorder
 - $\Delta S < 0$: decreasing disorder
- low pressure leads to higher entropy
- more bonds \rightarrow less chaos \rightarrow less entropy

gibbs free energy

- amount of energy in a system **available to do useful work** (g)
- $\Delta G = \Delta H \text{ (kJ/mol)} - T \text{ (K)} * \Delta S \text{ (J/mol*K)}$
 - 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**
- $-\Delta G \rightarrow$ free energy **decreases, spontaneous** reaction
- $+\Delta G \rightarrow$ free energy **decreases, spontaneous** reaction
- $0 \Delta G$: **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

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

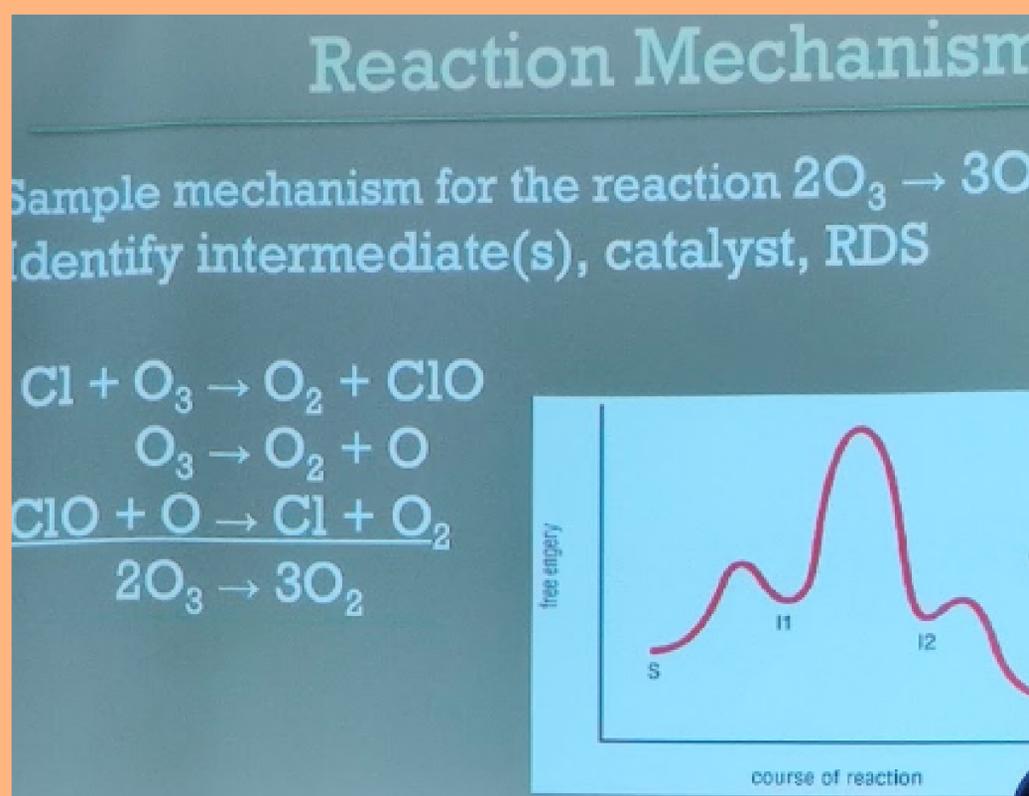


reactants: O_3 , product: O_2 , catalyst: Cl ,
intermediate: ClO

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**



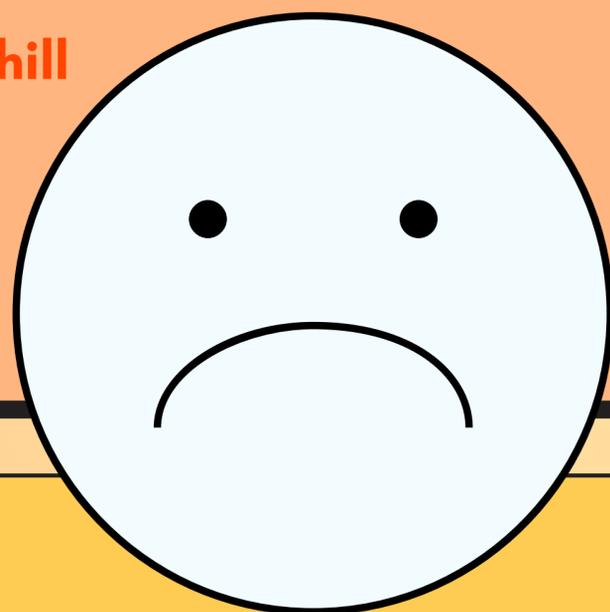
collision theory

- 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 (E_a): 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** E_a 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

