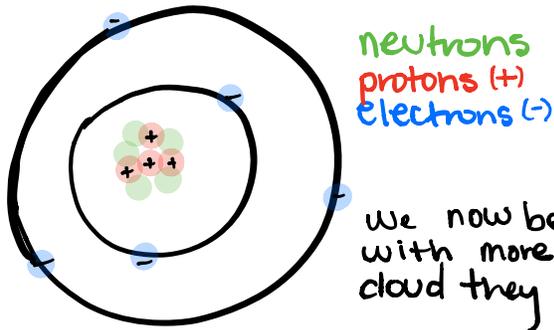
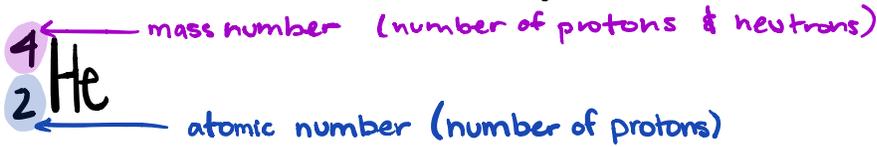


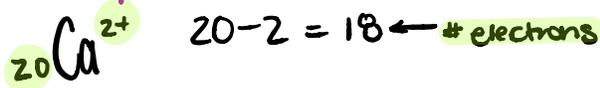
Bohr model of an atom



We now believe electrons orbit the nucleus with more of a wave pattern in a probable cloud they may exist in.



2 protons, 2 neutrons, 2 electrons



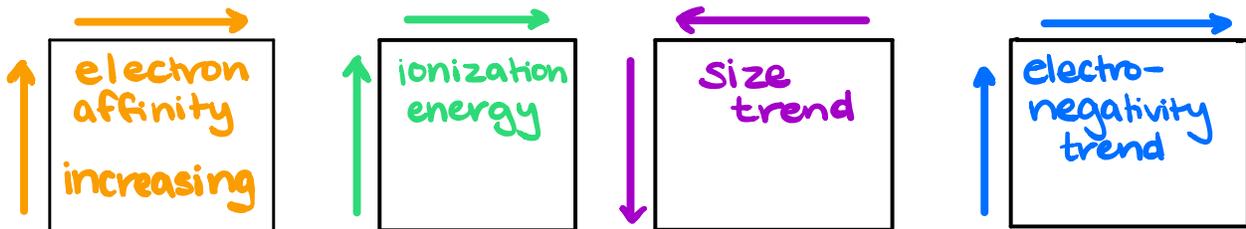
Isotope: ${}^4\text{C} \rightarrow {}^{12}\text{C}$
 ↑ ↑ Differing # of neutrons but same # protons

* Every element has multiple isotopes

 Cations
 * more protons than electrons
 (positive charge)

 Anions
 * more electrons than protons
 * larger than cations (they occupy a lot larger portion of empty space)

Periodic Table Trends



Periodic Table of the Elements

Group 1 → period

transition metals

alkali metals 1 valence e⁻ low ionization energy (reactive)

earth metals 2 ve⁻ higher ionization

Lanthanides

actinides

Halogen

Nobel Gas

1 H Hydrogen 1.01																	2 He Helium 4.00
3 Li Lithium 6.94	4 Be Beryllium 9.01											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.06	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.88	23 V Vanadium 50.94	24 Cr Chromium 51.99	25 Mn Manganese 54.94	26 Fe Iron 55.93	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.39	31 Ga Gallium 69.73	32 Ge Germanium 72.61	33 As Arsenic 74.92	34 Se Selenium 78.09	35 Br Bromine 79.90	36 Kr Krypton 84.80
37 Rb Rubidium 84.49	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium 98.91	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.6	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.91	56 Ba Barium 137.33	57-71 Lanthanides	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.85	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.20	83 Bi Bismuth 208.98	84 Po Polonium [208.98]	85 At Astatine 209.98	86 Rn Radon 222.02
87 Fr Francium 223.02	88 Ra Radium 226.03	89-103 Actinides	104 Rf Rutherfordium [261]	105 Db Dubnium [262]	106 Sg Seaborgium [266]	107 Bh Bohrium [264]	108 Hs Hassium [269]	109 Mt Meitnerium [268]	110 Ds Darmstadtium [269]	111 Rg Roentgenium [272]	112 Cn Copernicium [277]	113 Uut Ununtrium unknown	114 Fl Flerovium [289]	115 Uup Ununpentium unknown	116 Lv Livermorium [298]	117 Uus Ununseptium unknown	118 Uuo Ununoctium unknown
		Lanthanides										Halogen					
57 La Lanthanum 138.91	58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium 144.91	62 Sm Samarium 150.36	63 Eu Europium 151.97	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.97			
89 Ac Actinium 227.03	90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium 237.05	94 Pu Plutonium 244.06	95 Am Americium 243.06	96 Cm Curium 247.07	97 Bk Berkelium 247.07	98 Cf Californium 251.08	99 Es Einsteinium [254]	100 Fm Fermium 257.10	101 Md Mendelevium 258.10	102 No Nobelium 259.10	103 Lr Lawrencium [262]			
Alkali Metal	Alkaline Earth	Transition Metal	Basic Metal	Semimetal	Nonmetal	Halogen	Noble Gas	Lanthanide	Actinide								

= s block elements (2e⁻ max) ↓

= p block elements (have valence electrons in the p orbital)

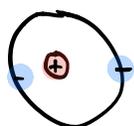
∞ 8 8 max 6 electrons

= d block: transition metals (10 electrons max)

* more unpaired e⁻ in d subshell = harder metal

= f block: Lanthanides (14 e⁻ max)

smaller atoms nuclei hold e⁻ tighter

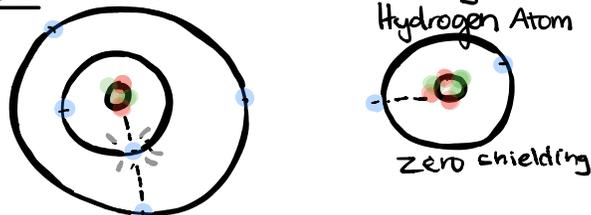


more (+) charge pulls (-) charge in tighter.

- more electronegative
- higher ionization energy

π bonds are formed by overlap of atomic orbitals. Large atoms don't overlap well so π bonds aren't formed.
poor orbital overlap.

Z effective : \uparrow atom size = \uparrow shielding



This pull is weaker than expected because inner e^- s shield the valence e^- .

Quantum Mechanics



1. "n" principle quantum number (shell)

Energy Levels =

- 1s
- 2s
- 2p
- 3s
- 3p
- 4s
- 3d
- 4p
- 5s
- 4d
- 5p
- 6s
- ⋮

2. "l" angular momentum quantum number

$0 = s$
 $1 = p$
 $2 = d$
 $3 = f$

} orbital

3. "m_l" magnetic quantum number

4. "m_s" electron spin quantum number (1/2)

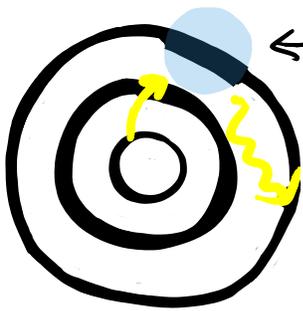
Electron Configuration

ex: $\uparrow \uparrow$
 \uparrow

anions: move forward 1 box per electron.
cations: move back 1 box per electron.

Energy Levels & Photon Emission

* Add energy to make e^- jump to a higher energy level. The e^- absorbs energy to jump. This is temporary until it falls back to the ground state. When it falls back to ground state it emits a photon (light).



← emitted photon

Energy of EP = difference in energy between the two levels.

$$E_3 - E_1 = E \text{ of photon}$$

$$E_{\text{photon}} = h\nu \leftarrow \text{frequency}$$

wavelength

$$c = \lambda \nu$$

↑ speed of light

↑ planks constant

$$* 6.626 \times 10^{-34} \frac{\text{m}^2 \text{kg}}{\text{s}}$$

If the photon that strikes the electron has a lower energy than $E_3 - E_1$ then it will not be absorbed.

$$* 3.0 \times 10^8 \text{ m/s}$$

$$\nu = \frac{c}{\lambda} \text{ or } E_{\text{photon}} = h * \frac{c}{\lambda} = E_3 - E_1$$

Radioactive Decay + 1/2 Life

↳ atom changes chemical composition over time.

α decay: Loss of 1 He nucleus

β decay: Neutron → proton (ejection of an electron)

Electron capture: proton → neutron (capture an electron)

positron emission: proton → neutron (expulsion of an electron)

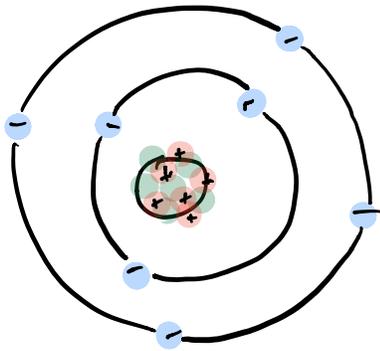
Gamma emission: Does not change # of nucleons (radiation from decay of nuclei)

α Particle: ${}^4_2\text{He}$
 ← mass
 ← charge

β Particle: ${}_{-1}^0e^-$ no mass, (-) charge

Positron: ${}_{+1}^0e^+$ absence of electron

γ Particle: ↑ λ high frequency, high energy



β decay: $\bullet \rightarrow \bullet + e^-$
 In β decay we eject an electron by a neutron converted to a proton

α decay: ${}^4_2\text{He}$ is lost

Electron capture: $\bullet + e^- = \bullet$

Positron emission: $\bullet = \bullet + e^+$
 expulsion of an electron

γ emission: Energy

Neutrons: $\bullet + \bullet = \bullet$

proton: $\bullet + \text{positron} = \bullet$

Bonding

1. Anti-bonding orbitals are higher in energy than bonding orbitals.
2. Bonding orbitals = e^- in phase (attractive)
3. anti-bonding orbitals = e^- out of phase (repulsive)

Ironic Bonds:
 • weaker bonds between ions



Covalent Bonds:
 • stronger bonds where ions are shared. No net dipole.



Properties

- Electrolytes: compounds that dissociate in water. (strong acids + strong bases)
- Ionic character: Difference in electronegativity (Francium - Fluoride would have the most ionic character)
- Conductivity:  concentration of NaCl to conduct electricity exactly as well as the solution.

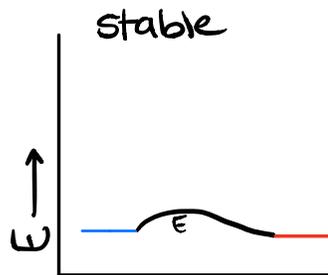
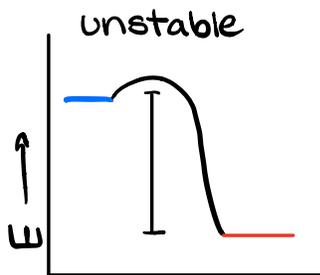
stable compounds (N_2) have high bond energy
 unstable compounds (ATP) have low bond energy

Bond Dissociation Energy = Bond Energy

Heat of Combustion: amount of energy released when a molecule is combusted with O_2 .

↑ Energy ↓ Stability ↑ heat of combustion = exothermic
 ↳ substance + O_2 = heat!

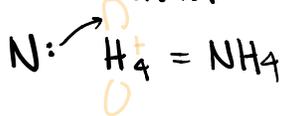
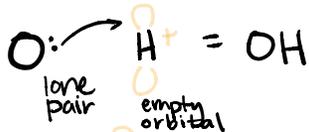
When a bond is formed... energy is released.
 When a bond is broken... energy is required.



The more unstable a bond the greater heat of combustion

Coordinate Covalent Bond

when electrons shared were all donated by one atom.



Empirical vs. Molecular Formula

↳ The most reduced form of the molecular formula.

↳ Tells how many of each atom is in a compound.

$$\% \text{ mass} = \frac{\text{mass of an element}}{\text{total mass of compound}} * 100$$

* Deriving a Formula from percent

1. Change % mass \rightarrow g (15% \rightarrow 15g)
2. Convert grams of each element into moles (divide by molar mass)
3. Look @ element with lowest # of moles. Get to all whole #'s.

You can only get Empirical Formula from % mass. You need compound MW to find molecular formula.

\rightarrow If you have a MW... divide that by the empirical formula MW to get a whole # and multiply by that.

EX

$$23.3\% \text{ Mg} = 23.3\text{g Mg} / 24.3 = .95 = 1$$

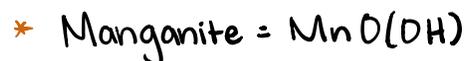
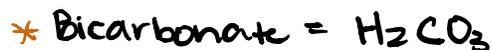
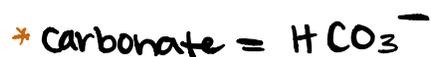
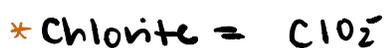
$$30.7\% \text{ S} = 30.7\text{g S} / 32 = .95 = 1$$

$$46.0\% \text{ O} = 46\text{g O} / 16 = 2.875 = 3$$



Inorganic Nomenclature

- Ionic compounds: name cation first, then anion
- Transition metals: In words, add the oxidation state in roman numerals (iron III sulfate).
- Monoatomic ions: add -ide (chloride = Cl^-).



Reactions

Combination Rxns: 2 or more reactants combine to form a single product.



Decomposition Rxns: A single compound breaks down into two or more elements or new compounds.



Single Displacement: An element reacts with a compound and takes the place of another element in the compound



Double Displacement: Two compounds react and the cations and anions switch places.



Balancing a Reaction

1. Balance carbons
2. Balance hydrogens
3. Balance oxygens
4. Balance other elements
5. Use fractions if necessary $7OH \rightarrow \frac{7}{2}O_2 + 7H$
6. Multiply everything by denominator of fraction
7. **Double Check Every Reaction**

Rxns for Combustion of methane, propane, propanol

1. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
2. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
3. $2C_3H_8O + 9O_2 \rightarrow 6CO_2 + 8H_2O$

Mole Stoichiometry

Atomic weight: mass of 1 mole of any atom
molecular weight: mass of 1 molecule (amu)
Molar mass: mass of a mole of molecules (g/mol)
mole: a unit of weight (6.022×10^{23})

Molar Conversions

moles \rightarrow grams = $\frac{g}{\text{mol}}$

$$\frac{15 \text{ grams H}_2\text{O}}{18 \text{ g}} \left| \frac{1 \text{ mol H}_2\text{O}}{18 \text{ g}} \right. = .833 \text{ moles H}_2\text{O}$$

\swarrow MW

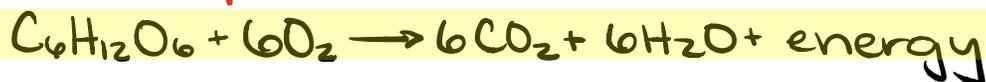
moles \rightarrow molecules

$$\frac{6 \text{ moles CaCl}_2}{1 \text{ mol CaCl}_2} \left| \frac{6.022 \times 10^{23}}{1 \text{ mol CaCl}_2} \right. = 3.61 \times 10^{23} \text{ molecules CaCl}_2$$

Limiting Reagent

1. Balance equation
2. Convert grams to moles
3. Compare # moles to moles required to run 1 reaction
4. The reactant you run out of first is the Limiting reagent

Respiration Reaction



Theoretical Yield: How much product we should produce based on calculations.

Actual Yield: How much product was produced in an experiment.

Percent Yield: The discrepancy between predicted and actual values.

$$\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 = \% \text{ error}$$

Percent Error: $\frac{\text{Theoretical} - \text{Actual}}{\text{Theoretical Yield}}$

••• Moles of oxygen needed to combust •••

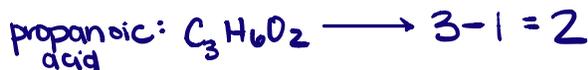
"predict the species that requires the most oxygen to combust"

I.e. which needs the most oxygen to completely burn up and be gone.

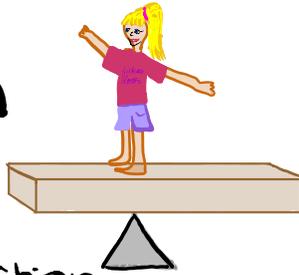
ADD 1 for every carbon.
SUBTRACT .5 for every oxygen.

The higher the number the more O₂ is needed for total combustion.

ex: 3 carbons



Equilibrium



The Law of mass Action

$$K_{eq} = \frac{[\text{products}]^x}{[\text{reactants}]^y}$$

coefficient in balanced equation

* Constant when @ Equilibrium
* reactants create products in same amount products create reactants.

progress of reaction

$$Q = \frac{[\text{products}]^x}{[\text{reactants}]^y}$$

@ Equilibrium

$$K = \frac{[\text{Products}]^x}{[\text{reactants}]^y}$$

↓
How fast a reaction reaches equilibrium

If Q < K then products are smaller than reactants so you need more products according to Le Chatliers.

If Q > K then products are larger than reactants

Q = K is equilibrium

* K changes with temperature.
@ 25°C K will be constant.

↑ Catalyst = ↑ Rate of reaction. (Constant K)

* ↑ Temperature = New K value.

↓ Energy of Activation = ↑ Rate to equilibrium (Constant K value)

Stabilized transition state = ↑ Rate of reaction (constant K)

↑ Reactants or products = ↓ Rate to equilibrium (constant K)

* The only thing that changes K .

Le Chatelier's Principle

Systems at equilibrium that experience change will shift to reestablish equilibrium.

- * Adding more reactants or product does not change K_{eq} .
- * Only temperature will change K_{eq} .

HOW TEMPERATURE SHIFTS A REACTION

For **exothermic** reactions: ↑ temperature shifts reaction to the left. (↓ K_{eq})



For **endothermic** reactions: ↑ temperature shifts reaction to the right (↑ K_{eq})



This reaction is giving off heat (meaning it is getting colder) and the equilibrium can't stay high so the K_{eq} decreases.



This reaction is taking in heat so there is more heat and adding heat causes equilibrium to increase. ↑ K_{eq} .

L R

Reactant → Product

↑ Products = Reaction shifts left

↑ Reactants = Reaction shifts right

↓ Products = Reaction shifts right

↓ Reactants = Reaction shifts left

↑ Pressure = Reaction shifts to side with fewer moles of gas } $PV = nRT$

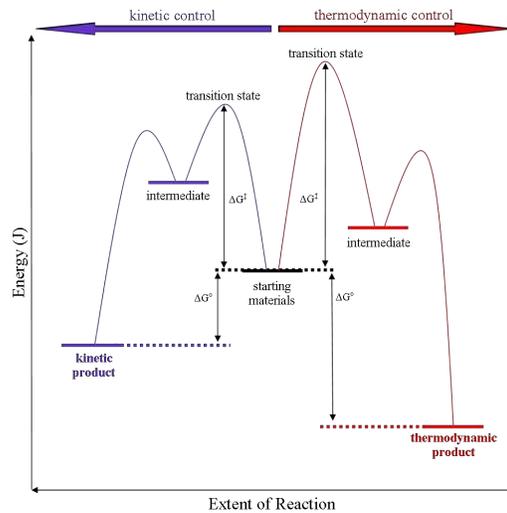
↓ Pressure = Reaction shifts to side with more moles of gas }

← Kinetics →

The energy of motion in a chemical reaction. i.e.: energy required to activate a reaction and it determines the reaction rate.

Thermodynamics: The gibbs free energy (ΔG).
*Nothing to do with rate!

Catalysts: Decrease the activation energy. And, decrease the kinetics.



Collisions cause reactions.

1. There is enough energy in the collision to overcome the energy of activation.
2. Reactants must be in correct spacial organization.

Rate change units: $\frac{\text{Molarity (M)}}{\text{second (s)}}$

Rate Laws

1. Write a rate law
2. Determine rate law from experimental values
3. Predict experiment from rate law

Assumptions

- * Rxns proceed forward
- * Consider first few rxn seconds
- * High concentration of reactant + catalysts.

Exponents: order of reaction

Calculating the order of each reactant

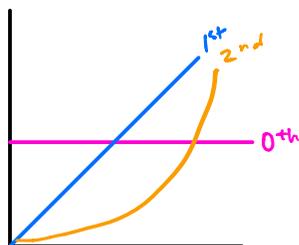
- 1) Find trial where [reactant] changed but other parameters didn't.
- 2) Note factor by which reactant concentration changed
- 3) Factor the rate that changed the same across two trials.
- 4) Solve for $Y: X^Z = Z$
X = factor that reactant changed
Z = factor that rate changed
Y = order of reactant

Overall order of reaction = sum of exponents



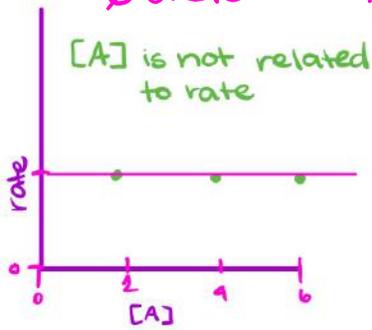
Rate order graphs:

- 0th order = [A] vs time is linear and slope = -k
- 1st order = $\ln[A]$ vs time is linear and slope = -k
- 2nd order = $\frac{1}{[A]}$ vs time is linear and slope = k

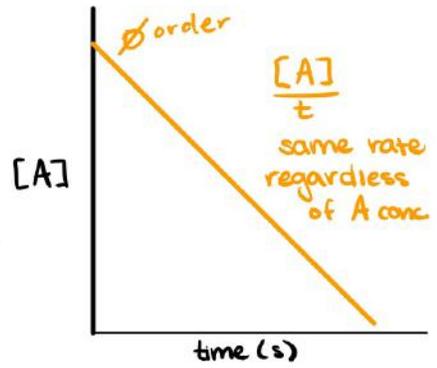


Rate Order graphs

Order: Rate = $k[A]^0$

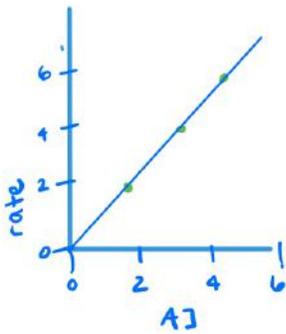


Rate	k	[A]	[A] ⁰
1	1	2	2 ⁰ =1
1	1	4	4 ⁰ =1
1	1	6	6 ⁰ =1

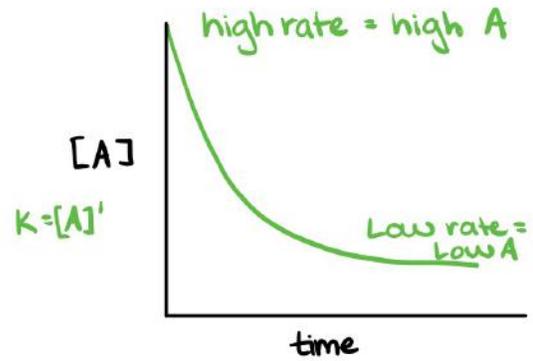


1st Order:

Rate = $k[A]^1$

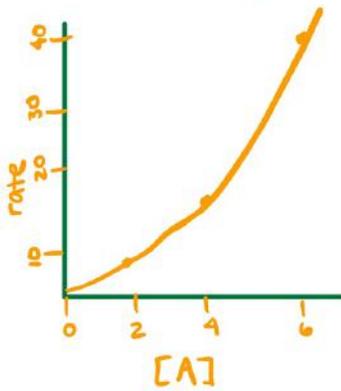


Rate	k	[A]	[A] ¹
2	1	2	2
4	1	4	4
6	1	6	6

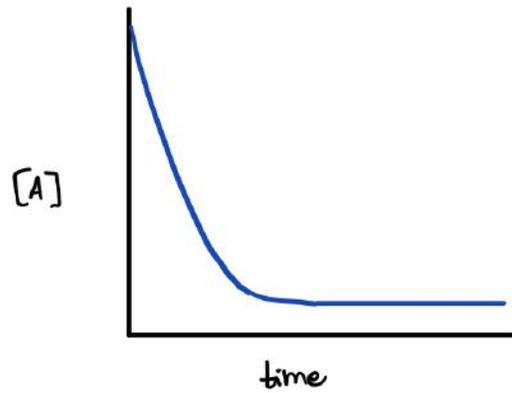


2nd Order

Rate = $k[A]^2$



Rate	k	[A]	[A] ²
4	1	2	2 ² =4
16	1	4	4 ² =16
36	1	6	6 ² =36



Heat Exchange

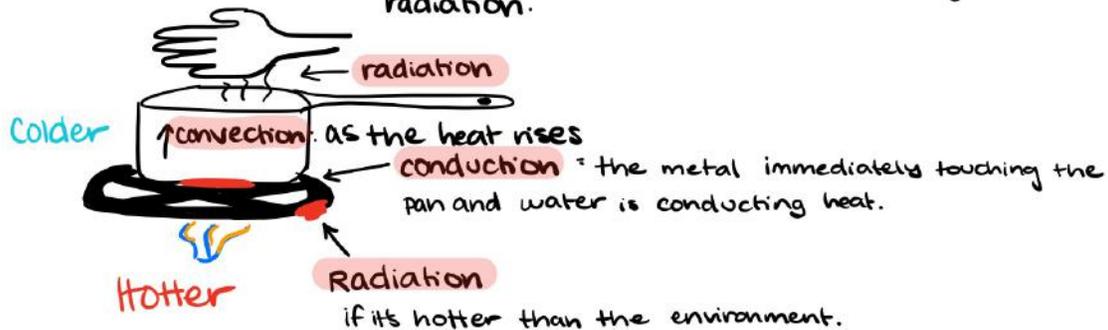
Heat exchange

conduction: molecule collision against something else
• ie touch a stove & it's hot & transfers heat

convection: hot air, less dense, goes up b/c weighs less than normal air

radiation: Electromagnetic field
~ Heat you feel from sun ~

Black Body Radiator: an object that absorbs all radiation falling on it, at all wavelengths. Then it emits 100% energy as electromagnetic radiation.



Heat Capacity = The energy to change the temperature

$$C = \frac{q}{\Delta T}$$



constant volume



constant pressure

constant volume = lower heat capacity
constant pressure = greater heat capacity

Temperature = average kinetic energy of molecules.

In a constant volume system all the heat flow is translated into heat (no work is done).

In a constant pressure system there is room for expansion so there is work done.

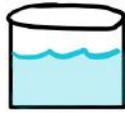
$$\Delta U = Q + W$$

Specific Heat capacity = "c" not "C"

$$q = mc\Delta T$$



.25L



.50L

- Which has the greater heat capacity?
- Which has the greatest specific heat capacity?

Heat capacity is higher for .5L.
Specific capacity is the same for both.

Calorimeters

Device used to calculate enthalpy change (ΔH)

atmosphere



coffee cup calorimeter

$$q = mc\Delta T$$

Constant Volume



Bomb calorimeter

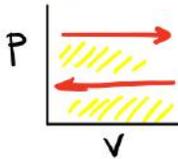
$$q = mc\Delta T = \Delta E$$

Constant Pressure

PV work

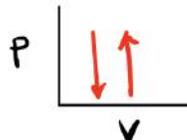
→ constant pressure. Change in volume.

* Look at area under the graph



= PV work

coffee cup calorimeter



= no PV work

Laws of Thermodynamics

1st Law of Thermodynamics: $\Delta E = \Delta Q + W$
↑
constant in isolated system

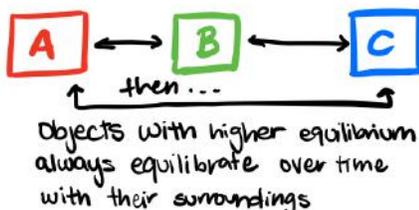
2nd Law of Thermodynamics:

- Entropy never decreases in an isolated system.
- Heat cannot entirely translate into work.

3rd Law of Thermodynamics:

- Crystalline substances at absolute zero have zero entropy. [metals + minerals]

0th Law of Thermodynamics:



Enthalpy ΔH : Energy contained in chemical bonds
measured in Joules

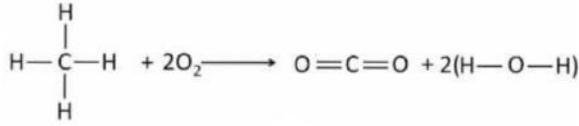
Standard state \neq standard temperature + pressure (STP)

↓
25°C = 298K

Elements in standard state = $\Delta H^{\circ}_{\text{formation}} = 0$
This is zero because standard state elements

If a bond is broken ... energy is required (+)
If a bond is formed ... energy is released (-)

Calculate ΔH for the following reaction, using average bond enthalpies, assuming all compounds are in their gaseous states :



Bonds Broken			Average Bond Enthalpies (kJ/mol)	
Type	Number	kJ/mol	Type	kJ/mol
C-H	4	$4 \times 413 = 1652$	C-H	413
O ₂			C-C	348
Bonds Formed			C-O	358
Type	Number	kJ/mol	C=O	799
			O-H	463
			O ₂	495



Balance :

	$\begin{array}{c} \text{H}4 \\ \text{C}1 \\ \text{O}4 \end{array}$		
Formed:	C=O x 2	- 799 x 2 =	-1598
	O-H x 4	- 463 x	= -1852
Broken:	C-H x 4	+ 413 x 4 =	1,652
	O-O x 2	+ 495 x 2 =	990
			-808 $\frac{\text{kJ}}{\text{mol}}$
			↑ exothermic

Entropy ΔS : measure of randomness or disorder in a system
 ↑ entropy = ↑ spontaneity

+ ΔS = more available energy
 - ΔS = less available energy

Joules · K

Equilibrium = max entropy

$\Delta S \uparrow$ by ...

- ↑ # particles
- ↑ volume
- ↑ temperature
- ↑ disorder
- ↑ complexity of molecule

Gibbs Free Energy : amount of energy free to do work

↑ Entropy = ↑ Free Energy (+)

↓ Enthalpy = ↑ Free Energy (-)

-ΔG = Spontaneous Rxn : has free energy

+ΔG = Non-Spontaneous Rxn

terms

Isobaric : Constant pressure process

Isothermal : Constant temperature process

adiabatic : Constant heat flow

$$\Delta G = \Delta H - T\Delta S$$

↑ ↑
ENTHALPY ENTROPY

*CHANGE IN BOND ENERGY

*MEASURE OF DISORDER

if: $K > 1$ then $G = -$

if: $K < 1$ then $G = +$

if: $K = 1$ then $G = 0$

Keq doesn't tell us anything about rate of rxn



(Exothermic = forming bonds)

Products



* higher bond energy
* greater enthalpy

Acid-Base Chemistry

$\text{pH} = -\log [\text{H}^+]$	$\text{pH} = \text{pKa} + \log \frac{[\text{A}^-]}{[\text{HA}]}$	$\text{pH} = \text{pKa} + \log \frac{[\text{B}]}{[\text{BH}^+]}$
----------------------------------	--	--

Acid dissociation = $\frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$ Base dissociation = $\frac{[\text{OH}^-][\text{HA}]}{[\text{A}^-]}$

* SMALL pKa = more products than reactants (↑ dissociation)
MORE Acidic

* LARGE Ka (Ka > 1) = MORE Acidic

$K_w = [\text{H}^+][\text{OH}^-] = 10^{-14}$
 $\text{p}K_w = \text{pH} + \text{pOH} = 14$

Strong Acids:

- * HI → Hydrogen Iodide
- * HBr → Hydrobromic acid
- * HCl → Hydrochloric acid
- * HNO₃ → Nitric acid
- * HClO₃ → Chloric acid
- * HClO₄ → Perchloric acid
- * H₂SO₄ → Sulfuric acid
- * H₃O⁺ → Hydronium

Strong Bases

- * NaOH
 - * KOH
 - * LiOH
 - * NH₂⁻
 - * H⁻
 - * Ca(OH)₂
 - * Sr(OH)₂
 - * Ba(OH)₂
 - * Na₂O
 - * CaO
- } Group 1A Hydroxides

Question?

What is the pH of a 1×10^{-3} M HCl solution?

shortcut

add decimal
 $3 - .1 \approx 2.9$ or $\approx \boxed{\text{pH} = 3}$

Strong acid

What is the pH of a 2×10^{-5} M HBr solution?

shortcut

$5 - .2 = 4.8 \approx \boxed{\text{pH} = 4.8}$

Strong acid

Calculating the pH of a weak acid:

What is the pH of a .2M HF solution? $K_a = 3.4 \times 10^{-4}$

$$K_a = \frac{[H^+][F^-]}{[HF]} \quad \text{replace with "x"}$$

$$\frac{(x)(x)}{.2-x} = 3.4 \times 10^{-4}$$

weak acid

omit this x because it is easier and will not change the answer significantly.

Do this math using scientific notation

$$\begin{aligned} x^2 &= (.2) * 3.4 \times 10^{-4} \\ x^2 &= 6.8 \times 10^{-5} \\ x &= \sqrt{6.8 \times 10^{-5}} = 8.24 \times 10^{-3} \end{aligned}$$

shortcut

$$3 - .824 = \approx \boxed{2.18 \text{ pH}}$$

Salt of weak Acids

HCO_3^- = weak acid

↳ CO_3^{2-} = conjugate base

↳ Na_2CO_3 = "salt" of weak acid

Salt of weak Base

NH_3 = weak base

↳ NH_4^+ = conjugate acid

↳ NH_4Cl = "salt" of weak base

Question: Which salt will increase the pH? (aka form a W.B.)

- NaNO_2
- NH_4Cl
- NaF
- NaClO_2
- CH_3COONa
- NaCl

* Conjugate of NO_3^-

ANSWERS

* F^- conjugate of HF

* ClO_2^- conj of HClO_3

* CH_3COOH

* HCl

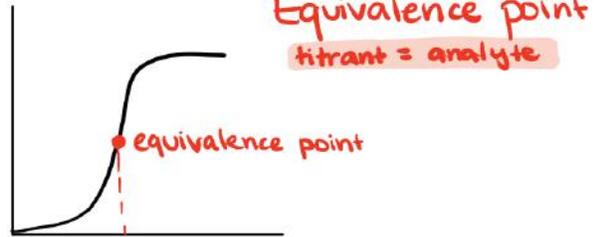
Titration Curves

The strong base is in the beaker as the analyte.
The strong acid is added drop by drop as the titrant.

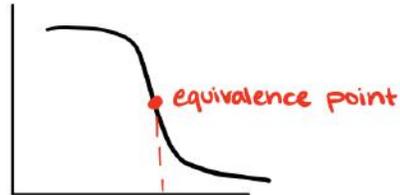
One equivalent = how much acid or base is needed to produce one mole of $[H^+]$ ions.

Titration Curves

Strong Acid titrated with Strong Base



Strong Base titrated with Strong Acid



Weak acid titrated with Strong Base



Weak Base titrated with Strong Acid



for Strong acids + Strong bases...
they completely dissociate in water 100%
so...
 $[H^+] = [OH^-]$ @ equivalence point
and $pH = 7$

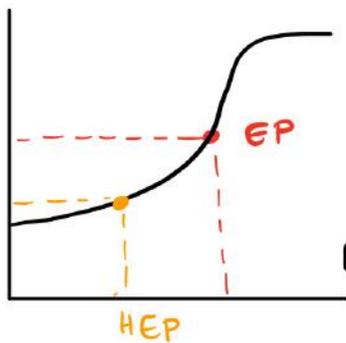
for weak bases + SA or weak acids + SB then $[H^+] \neq [OH^-]$

- Weak base w/SA = $pH < 7$
- Weak acid w/SB = $pH > 7$
- Strong acid w/strong Base = $pH \approx 7$

MCAT Q's :

What is the pH of the solution at the equivalence point?

HALF EQUIVALENCE POINT



↳ is where $pH = pKa$
and $[HA] = [A^-]$

* SA/SB titrations do not have half equivalence points

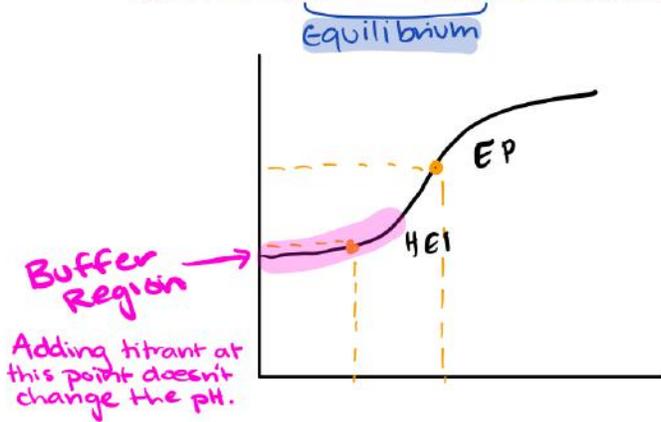
$[OH^-] = [H^+]$ (only in SA/SB @ EP)

Indicators: weak acids that change color as they dissociate from HA into H^+ and A^- .

End point: when the indicator causes the color change.

Buffers:

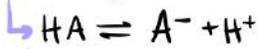
weak acid + weak base



Henderson-Hasselbalch:
$$\text{pH} = \text{pKa} + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

Half Equivalence

$\text{pH} = \text{pKa}$ @ half equivalence point



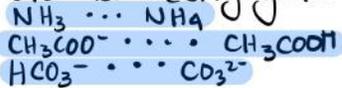
$\text{pH} = \text{pKa} + \log \frac{[\text{A}^-]}{[\text{HA}]}$

$\text{HA} = \text{A}^-$ at the half equivalence point so $\frac{\text{A}^-}{\text{HA}} = 1$ and $\log(1) = 0$ so pKa and pH are equal

MCAT buffer problems:

- Watch for equimolar amounts where weak acids and conjugate bases are equal

- Watch for conjugates...



- Weak acids + bases **STRONG ACIDS/BASES** \neq buffers

A SB or SA would immediately and dramatically change the pH because there would be no weak acid or base to consume the added buffer.

- watch for resistance to pH change
- watch for half equivalence point
- watch for the $\text{pH} = \text{pKa}$
- watch for the $\frac{[\text{A}^-]}{[\text{HA}]}$ ratio

Electrochemistry

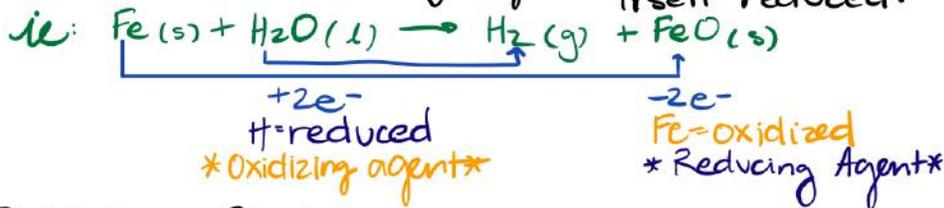
Redox Reactions

Oxidized = losing electrons

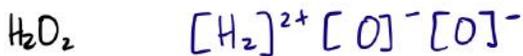
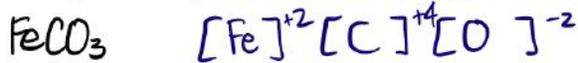
Reduced = gaining electrons

Reducing agent = donates electrons and is itself oxidized in the process.

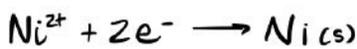
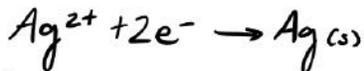
Oxidizing agent = accepts electrons and is itself reduced.



Oxidation State



Electrical Potentials



$E^\circ = .80\text{V}$

$E^\circ = .52\text{V}$

$E^\circ = -.23\text{V}$

$E^\circ = -.76\text{V}$

$E^\circ = 0.00\text{V}$

Hydrogen
Half
Cell
(Reference)

* the degree to which these species want electrons. → Reduction Potential *
Positive E = more likely to gain e^-
negative E = less likely to gain e^-

Solid metals do not get reduced

Cell Potential

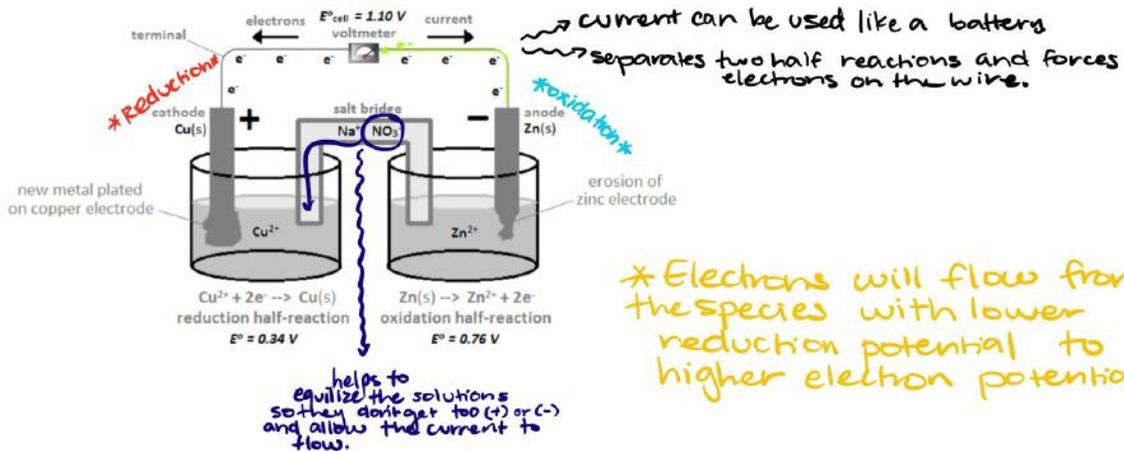
E°_{cell} is the sum of electric potentials for two half reactions.

↳ You must reverse the half reaction with the lowest reduction potential.

The Galvanic Cell

Cell potential = POSITIVE

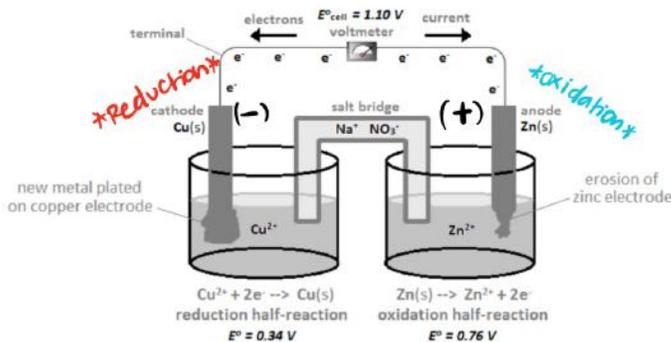
Chemical energy → electrical energy



The Electrolytic Cell

galvanic cell + external voltage which forces electrons in opposite direction

Cell Potential = NEGATIVE



Q: Species X has a reduction potential of .88V. Species Y has a reduction potential of .23V... if an electrolytic cell is constructed using these two metals, which metal will be at the cathode?

Species Y

Concentration Cell \rightsquigarrow type of Galvanic cell but the same electrodes and solutions are used in both beakers

E° cell of a concentration cell = 0.00V

$$E = E^\circ - \left(\frac{0.06}{n}\right) * \log \frac{[\text{lower}]}{[\text{higher}]}$$

↑
moles of e^- transferred

} Nernst Eqn. \rightarrow calculate cell potential based off E° of the species + solutions concentrations

Free Energy & Chemical Energy

$$\Delta G^\circ = -nFE^\circ \rightarrow +E^\circ = -\Delta G = \text{spontaneous Rxn}$$

↑ ↑
Faraday's constant (charge on 1 mole e^-)
moles e^- transferred

$$1e = 1.6 \times 10^{-19} \text{ Coulombs}$$

$$6.022e^{23} = 1 \text{ mol}$$

$$6.022e^{23} * 1.6 \times 10^{-19} =$$

$$9.6352 \times 10^4 \text{ C/mol}$$

$$PV = nRT$$

$$.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol}} * \text{K} \quad \text{OR} \quad 8.314 \text{ J/mol} \cdot \text{K}$$

1) Negligible of volume

2) IMF are negligible

The combined gas law:

$$\bullet \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

• Boyles Law
 $P_1 V_1 = P_2 V_2$

(constant temperature)

• Charles Law
 $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

(constant pressure)

STP

Standard temp & pressure

$$P = 1 \text{ atm}$$

$$V = 22.4 \text{ L}$$

$$n = 1 \text{ mole}$$

$$R = .0821 \text{ or } 8.31 \text{ J/mol} \cdot \text{K}$$

$$T = 273 \text{ K}$$

Absolute zero = 0K

\rightarrow all molecular motion ceases

*** partial pressure *** = $P_{tot} = P_1 + P_2 + P_3 \dots$

diffusion & effusion

gas molecules spread from high to low concentrations due to random kinetic energy and collision

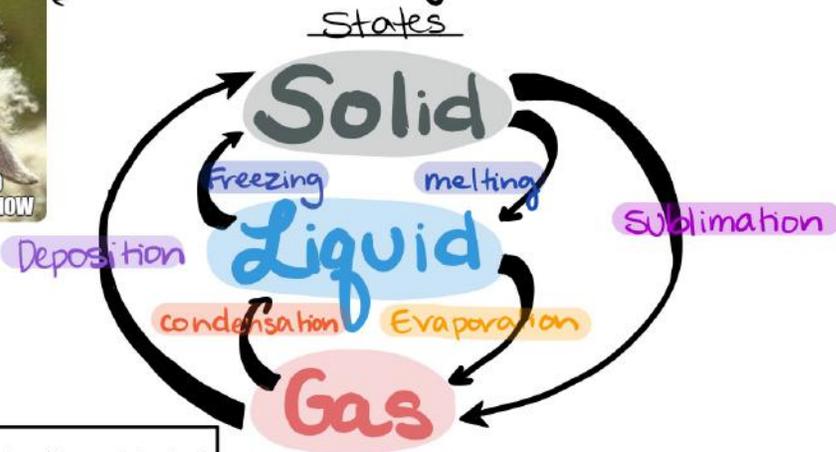
diffusion of gas through a pin hole.

$$\frac{E_1}{E_2} = \frac{\sqrt{MW_2}}{\sqrt{MW_1}}$$

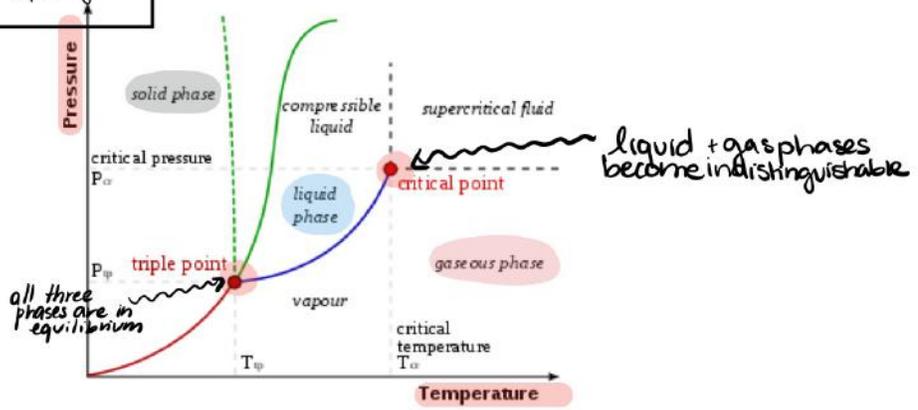
rate of effusion or diffusion is inversely proportional to molecular weight of the gas



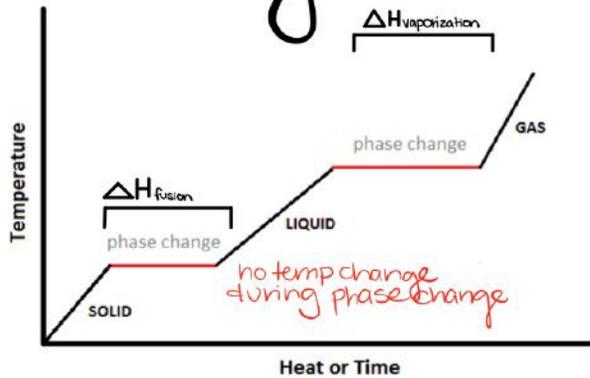
Phase Changes



ΔH_{fusion} : Energy required to go from solid \rightarrow liquid or liquid \rightarrow solid
 $\Delta H_{vaporization}$: Energy to go from liquid \rightarrow gas or gas \rightarrow liquid.



Heating Curve



Q: For water will the horizontal be longer between solid and liquid or between liquid and gas?

A: Between liquid and gas because more intermolecular forces must be broken. (all H-bonds must break to be a vapor.)

Once a phase change starts... all energy goes towards breaking intermolecular forces, none goes towards an increase in temperature

vapor pressure {
 ↑ temperature = ↑ vapor pressure
 ↑ non-volatile solute = ↓ vapor pressure
 ↑ volatile solute = ↓ vapor pressure

atm. pressure = vapor pressure
 @ boiling point

Raoult's Law

$$V_p = X V_p^{\circ} \quad (\text{pure solvent})$$

$$V_p = V_{p \text{ solvent}} + V_{p \text{ solute}}$$

gas solubility {
 ↑ temp = ↓ solubility
 ↓ temp = ↑ solubility
 ↑ vapor pressure = ↑ solubility of gas in liquid

OSMOTIC PRESSURE

$$\Pi = i M R T$$

↑ #ions ↑ solute molarity ← gas constant ← temp

* The side that receives the water = higher osmotic pressure

↳ ↑ solute = ↑ osmotic pressure

Two ideal solutions are separated by a synthetic biological membrane which mimics plasma membrane physiology. As the solutions proceed toward equilibrium which direction will water flow?



Solution A
 100 mM Na_3PO_4 \rightarrow Na Na Na PO_4
 Solution B
 200 mM NaCl \rightarrow Na Cl
 Semi-Permeable Membrane

^{ions}
 $4 \times 100 = 400$
 $2 \times 200 = 400$
same so

No net movement of water will occur.

ALWAYS SOLUBLE:

nitrate, ammonium, all alkali metals

insoluble:
 unless paired w/
 "always soluble"

carbonate, phosphate, silver (Ag), mercury (Hg) and lead (Pb).