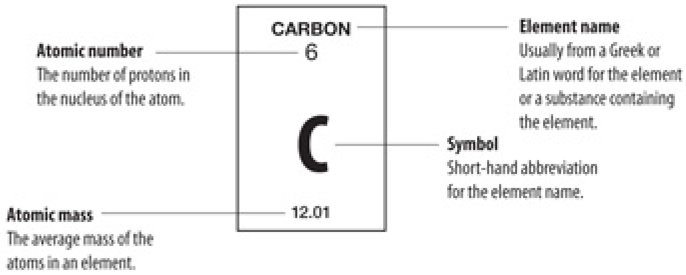
Get that 5 fellas

Due to too many people who failed the test coming on here and adding pointless edits that ruin the sheet, suggestions and edits are turned off.
Thank you to everyone who suggested good things and gave helpful edits. I hope this paper is as much of a help to you as it was to me.

Periodic Table

Atom- Smallest form of matter. Over 99% of the mass is composed of protons and neutrons in the nucleus, which is orbited by electrons.
Element- Specific configuration of an Atom
Compound- specific configuration of elements, whole number ratios

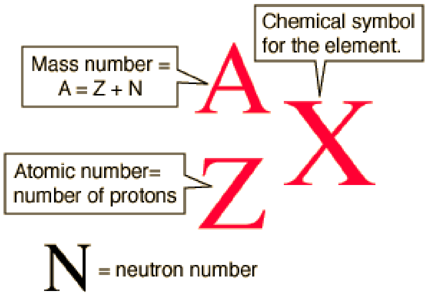


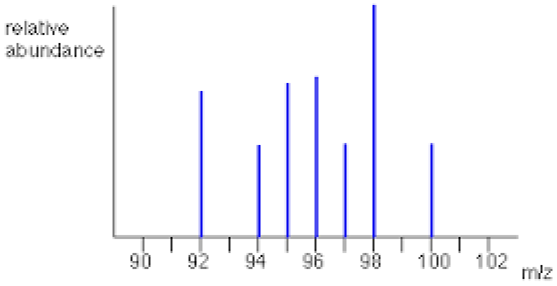
Horizontal Rows = Periods
Vertical Rows = Groups

Groups:
Group IA/1 - Alkali Metals
Group IIA/2- Alkaline Earth Metals
Group B/3 - Transition Metals
Group VII/17 - Halogens
Group VIII/18 - Noble Gases

Atomic Mass - Atomic Number = Number of Neutrons
Different Number of Neutrons are called Isotopes
Carbon-12 is normal
Carbon-14 is an isotope of 6 protons and 8 neutrons

Isotope Notation:





Element X is normally X-98
Element X could be Isotopes in 92, 94, 96, etc.

Average Atomic Mass = (Mass Isotope 1*%Abundance Isotope 1) + (Mass Isotope 2*j %Abundance Isotope 2) + (Mass Isotope 3*%Abundance Isotope 3) + etc...etc...

Valence electrons

Number of Valence Electrons = Main group number - (charges)(electron number)
Electrons = -1 charge
Example: N 2+ = 5 (main group number) - 2 (2 charge electrons) = 3 valence electrons

MOLES

Avogadro's Number : Defined constant for the amount of molecules present in one mole.
(6.022 x10^23)

Convert Moles to Grams (vice versa)
Moles = (gram/molar mass)
Grams = (Moles x Molar Mass)

Moles of a gas are found with n=PV/RT (Rearranged Ideal Gas Law.)
Most likely will occur at STP (Standard Temperature & Pressure)

Pressure = 1 atm
Temperature = 273 K (0 C)
At STP one mole of an ideal gas will be at 22.4 Liters.

P = Pressure in atm
V = Volume in liters
n = # of moles
R = Gas Constant (.0821 L*atm/mol*K)
T = Temperature in Kelvin (Take Celsius and add 273 to get Kelvin)(Test won’t give Fahrenheit)

Moles = (molarity)(liters of solute)

Percent Composition- percent of mass of each element in a compound
Divide each element's molar mass by compounds molar mass
Ca(NO3)2 = 164.10
Calcium = 1 = 40.08 x 1 = 40.08
Nitrogen = 2 = 14.01 x 2 = 28.02
Oxygen = 6 = 16 x 6 = 96.00
Ca = 40.08/164.10 = 24.42%
N = 28.02/164.10 = 17.07%
O = 96/164.10 = 58.50%

Empirical Formula

Empirical formula can be found by calculating the moles of each element presented, and then dividing all mole values by the smallest number found. The values given by the division will be the empirical formula subscripts.

Electron Configurations

Nucleus is positive, pulling in electrons which are negative, closer they are stronger the attraction is i.e More protons= more attraction, closer electrons= more attraction

Electrons are repelled by other electrons, an electron between a valence and nucleus causes the valence to be weaker (called shielding)

Completed shells are stable, atoms strive for completed shells

Aufbau principle- when building electron configurations, electrons go into shells in order of increasing energy

Pauli Exclusion Principle- two electrons in a shell will not spin the same way

Hund's rule- when an electron is added it will take up the lowest energy orbital

Amount of energy from an electron is dependent on how far away it is from the nucleus

Bohr Model - Quantum model for the atom where electrons are arranged in specific orbits around the atom - used this to predict the emission spectrum of hydrogen given the electron distribution. Problem - good for predicting some emission spectrums but does not work with most. Also, electrons do not "orbit" in a perfect, flat plane.

If an atom is exposed to electromagnetic levels greater than the ones exerted by the electrons it holds the electrons can be ejected.

Electron Configuration w/ periodic table

Use Periods (horizontal) as first number, then letter of the group, then amount
Examples
    Hydrogen: 1s^1
    Helium: 1s^2
    Lithium: 1s^2, 2s^1
    Sodium: 1s^2, 2s^2 2p^6, 3s^1
You can also 'shorthand' by using the Noble gasses as starting points
    Argon: [Ne] 3s^2, 3p^6

Moving left to right from periodic table atomic radius (distance from nucleus to valence electrons) decreases (Zeff increases while valence electrons are added to same energy level, so overall the electrons are pulled in more)

Moving down a group, atomic radius increases (more shells mean the valence electrons are further out)
X

Removing electron creates a positively charged cation (X1+), the energy to remove the first electron from a gaseous ion is the ionization energy, the energy to remove the next electron is the second ionization energy, which is always higher than the first ionization energy as the number of protons remains the same while there is one less electron, so less force repelling and the same force attracting, overall requiring more force to remove the electron.

If the electron being removed drops a shell, ionization energy increases MASSIVELY.

Electronegativity
- Small atoms have larger electronegativity levels
- Atoms close to having stable electron amounts have stronger electronegativity levels

BONDS

Bonds - sharing of electrons, more electrons shared = stronger bond
So single bonds (1 σ) have the least bond energy
Bond Energy Strength: triple (2 π and 1 σ)>double (1 π and 1 σ)>single (1 σ)

Ionic Bond- Bonds held by electrostatic forces, one will be + (cation(red)) while the other is - (anion(black))
Ionic bonds are where the anion completely removes the electron from the cation
Strength of bond depends on size; smaller size=greater attraction

Metallic Bond- Metals in the bond are usually stationary and share a sea of electrons, making them good conductors as they give up electrons easily,this also makes the metals malleable

Metals can bond with each other and form alloys
- Interstitial Alloy- metals with VERY different atomic radii combine
  - Example is carbon filling in the cracks around iron to make steel

Substitutional Alloy- Two metals with similar atomic radii combine
Example is Zinc and Copper, it is much more structured and uniform

Covalent Bonds- Two atoms share electrons and both count the electrons as part of their valence shell (nonmetals)

The first covalent bond is called the sigma bond (σ), and has the lowest bond energy and longest bond length, while the following bonds (2nd and 3rd) are called pi bonds

Network Covalent Bonds-these bonds are formed when a group of atoms share covalent bonds into one big electron sharing cluster, it is very hard to break and has a high melting point

Good contenders are **Silicon** and **Carbon** since they only have 4 valence electrons (ex. SiO_2) <--saw this as a suggestion from SiO_4, its been years since i did chem which is correct?

Polarity- Atoms usually will have an electronic imbalance of sorts due to uneven distribution of electrons (very based on symmetry)
Dipole- The most electronegative element will pull on electrons but not enough to become ionic, but will keep them on one side of the atom, H_2O for instance has 2 H^+ on one side, connected to O^- this makes the O^- side negative, and the H^+ side positive, more charges means more dipole moment, which is how a dipole is measured
Dipole moments can attract each other, causing the H^+'s in H_2O to be attracted (they do not form a bond) to the O^-'s in other water molecules (this is a hydrogen bond described below)

Intermolecular forces- forces that exist between 2 or more covalent molecules that hold them together. (ex. London Dispersion forces, dipole/hydrogen bonds.) Stronger IMFs = higher boiling point, lower vapor pressure

Hydrogen Bonds- Special type of dipole moment attraction that involves a positively charged hydrogen to bond to a negative end of another molecule, usually one with high electronegativity (H—F, H—O, H—N).
Extremely strong bonds, such as hydrogen, give up it's only electron so it uses all its strength to hold onto it, molecules with these bonds have high boiling/melting points. A group of H-F molecules, or any molecule w hydrogen bonding,are oriented so that the very positive H of one molecule is attracted to the very electronegative F of another molecule.

London Dispersion forces- With large amounts of electrons moving in random patterns it is possible for a majority to move to a specific side, creating a quick bit of polarity (Dipole Moment), which can in turn disrupt other molecules. More electrons means more chance of them congregating (more polarizable), so molecules with high boiling points are the most vulnerable. Larger molecules are more polarizable so have more electrons and will more likely have London dispersion forces
To compare the forces, compare the IMFs - larger is stronger

Bond order = (# of sigma and pi bonds/ # of atoms)
Formal Charge = Group Number - Lone electrons - # of sigma and pi bonds
(Group Number - dots - sticks = formal charge)

Bond Strength + Intermolecular Forces (IMFs)
Strongest to weakest

Covalent Bonds

(Usually liquids.)

Network Covalent Bonds are strongest possible (think diamond).
(sharing valence electrons means you need to break 2 sets of valence electrons, making it the toughest to break)
(Melting/boiling point is lower than ionic)
Use 4,5,6 bonds below to base on other covalent bonds strength
Electrons are shared.

Nonmetal-Nonmetal.

Ionic Bonds
Usually solid.
Energy to break based on coulombic attraction
Bad conductors in solid form as electrons don't move much.
Generally good conductors in aqueous solution.
Electrons are transferred.
Metal-Nonmetal.

Metallic Bonds
Transition metals seem to be strongest of all metals
High melting point, usually below ionic
Generally good conductors as a solid. ('Mobile sea of electrons.')
Metal only.

Hydrogen Bonds (Actually an IMF despite 'bond' being in the name.)
Special/specific kind of Dipole attraction
Most commonly between H and F,O,or N ("Hydrogen bonding is FON!")- corresponding F,O, or N must have a pair of unpaired e-
Strongest IMF.
Contains both LDFs and DD.

Non-Hydrogen Dipole bonds (Dipole-Dipole - DD)
Usually long lasting
Stronger than LDFs but weaker than HB.
Contains LDFs
Polar

London Dispersion Forces (Sometimes called Induced Dipole)
Only creates a dipole temporarily
Weakest IMF.
Is the reason why larger molecules tend to have higher boiling points/lower vapor pressure than smaller ones, can be used to help determine which molecules have the stronger IMFs if they

both have D-D and/or HB.
Greater # of electrons corresponds to greater polarizability, which makes for stronger LDFs
Every molecule has LDF forces regardless of if they have other IMFs, be it dipole dipole or H bonding.

Molecules with weak IMFs (4,5,6 above.) tend to be gases at room temp (Generally have ONLY LDFs)
Molecules with stronger intermolecular forces tend to be liquid at room temp. (Strong LDFs, DD, HB.)
Ionic substances do not have Intermolecular forces. (Usually solid at room temp)
If a molecule in a liquid builds up enough energy it can break free into a vapor, this is vaporization. (This is also responsible for vapor pressure.)
Single < Double < Triple bonds in bond strength
Single > Double > Triple in bond length

Lewis Structures + VSEPR:

It is important to remember that lewis structures are a visual representation of bonds and electrons.

For single atoms/ions all valence electrons should be drawn in. It is important to note that if something is ionized and has no electrons left, then no dots are drawn in - do not go down to the next shell! Also, ions should be put inside of brackets with the charge indicated in the upper right corner.

In molecules, bonds are represented by drawing a line between the two atoms. These lines can be used to represent single, double or triple bonds between two atoms in a molecule.

An important thing to remember is that each bond represents two electrons, any unbonded electrons should be shown in pairs (two dots) (with the exception of an odd total for valence electrons) and that the total number of electrons should be equal to that of the sum of the valence electrons of the atoms. (Make sure to add/subtract the charge from this total if working with polyatomic ions!) (Negative adds electrons, positive subtracts electrons!)

Lastly, as with single atoms, molecules should be drawn inside of brackets if they are ions.

VSEPR (Valence Shell Electron Pair Repulsion) Theory : A model used to predict the geometry (structure) of molecules based on bonds and lone pairs of electrons. The idea is that electron pairs will arrange themselves in a way so as to minimize the repulsion between them.

Electron Domain Geometry : The shape of the molecule that arises from the configuration of the molecule about the central atom. The geometry is determined by the number of electron domains (Also called Steric Number.) regardless of whether they are bonds or lone pairs of electrons. Often referred to as hybridization (only sigma bonds hybridize).

Molecular Geometry : A more specific shape of the molecule that arises from the configuration of the molecule about the central atom. The geometry is determined by the number of bonds as opposed to the total number of electron domains.

*****BOND ANGLES ARE DETERMINED BY THE ELECTRON DOMAIN GEOMETRY!*****

REACTIONS

Assume molecules without (solid, liquid, or gas) are aqueous (dissolve in water)

Synthesis- When two simple compounds form to make a more complex one
$2H_2 + O_2 \rightarrow 2H_2O$ (liquid)

Decomposition- opposite of Synthesis, complex compound breaking into simple compounds, usually requires energy (heat), usually endothermic

$2MgO \rightarrow 2Mg + O_2$

f-Base- when an acid(H^+) reacts with a base(OH^-) to form a salt and water

$HCl + NaOH \rightarrow NaCl + H_2O$(liquid)

Oxidation-Reduction- reaction that changes the phase of a species (Covered in greater detail later on) (involves transfer of electrons, used in batteries)

$Cu^+ + e^- \rightarrow Cu$(solid)

Precipitation - Two aqueous solutions mix, sometimes a precipitate (solid excess compound) is formed, also a cation and anion mixing can create a salt precipitate (non aqueous)

$K_2CO_3 + Mg(NO_3)_2 \rightarrow 2(KNO_3) + MgCO_3$(solid)

Also since the $K+$ and NO_3- both start out aqueous and end aqueous, they don't really do much and are called spectator ions

Net ionic equation - break up the molecules (ex. K_2CO_3 into $2K^+ + CO_3^{2-}$), then remove all spectator ions from the equation

The AP test will provide you with most solubility rules, the ones it will not tell you are:
Compounds with (positive) Alkali metals are always soluble (Li^+, Na^+, K^+, Rb^+, Cs^+, Fr^+)
Compounds with NO_3^- are always soluble.
Compounds with NH_4 are always soluble.
Chloride (As well as Bromide and Iodide) compounds are usually soluble, the exceptions being when paired with Ag^+, Hg^+, and Pb^+
Compounds with SO_4^{-2} are usually soluble, the exceptions being when paired with Ba^{+2}, Sr^{+2}, Hg^{+2}, and Pb^{+2}
Hydroxides are insoluble except when in a compound with any of the first three bullets.

BALANCING CHEMICAL EQUATIONS

Reactant side (starting compounds go here)→ Product side (compounds after reaction go here)

When balancing an equation, all elements or compounds must be equal (Conservation of mass)
For example, $2(H_2O) \rightarrow 10(H_2) + 5(O_2)$ is incorrect, because the reactant side only has 4 hydrogens, and 2 oxygens

However $2(H_2O) \rightarrow 2(H_2) + O_2$

REDOX

For a reaction to be a REDOX species in the reaction must undergo a change in charge.
Reduction is when electrons are gained.
Reduction is Gain: RIG
Gaining Electrons Reducing: GER
Oxidation is when electrons are lost.
Oxidation is Loss: OIL
Losing Electrons Oxidation: LEO
Just remember OIL RIG/LEO goes GER
If an atom is by itself on one side of the reaction, and then on the other side of the reaction it is in a compound that should be an instant clue to you that this is a REDOX reaction, because that means that the charge changes
Some Important rules for determining charge to follow are:
An element by itself always has a charge of 0 (Even diatomics!)
Oxygen will always have a charge of -2 (Unless by itself or if it is in the peroxide molecule O_2^{-2}

in which case each oxygen will have a charge of -1.)
Hydrogen will have a charge of +1. (Unless by itself, or in the case that it has a -1 charge - don't worry about remembering this, it won't happen unless they tell you it will.)
Group 1 and 2 elements usually have a +1 and +2 charge, respectively.
Group 17 elements will usually take on a -1 charge.
While group 1, 2 and 17 elements usually only have one oxidation state, transition metals can have several.
The overall charge of the compound should be zero unless it is an ion, in which case it should then add up to the indicated charge of the ion. (Unless it is a polyatomic you should have memorized, in which case you should have also memorized the charge - it may not be given to you.)
Ex :

Reaction : AB + C A + CB

Let's assume 'A' and 'C' are group one elements and 'B' is a group 7 element. On the reactant side A has a charge of +1, B has a charge of -1, and C is neutral. We know this because A is a group 1, B is a group 17 and C is by itself. On the product side, A now has a charge of 0, B still has a charge of -1, and C now has a charge of +1. We know this because A is by itself, C is a group 1 element, and B is a group 17 element.
This means that:

A has gone from a charge of +1 to 0. It has been reduced, or has gained electrons/become more negative. (why can't electrons be positive reeeeee)

B has not undergone a change in charge - no change means it is not part of the redox! (This is important in Net Ionic and Half Reactions, defined and explained below.)

C has gone from a charge of 0 to +1. It has been oxidized, or has lost electrons/become more positive.

Balancing REDOX:

Usually the types of REDOX reactions we should expect to deal with are not so simple as the example given above. They generally will require balancing the equation as well, and when this is the case, they usually also need to be broken down into half reactions so as to maintain both a conservation of charge and mass.
Full Ionic Equations represent all species present in the reaction, whether they 'participate' or not. Weak electrolytes are NOT split up into ions. (These are also good to use when determining precipitates if you are shaky on solubility rules.)
Example :
HF(aq) + NaOH(aq)NaF(aq) + H2O(l)* will become
HF(aq) + Na+(aq) + OH-(aq)Na+(aq) + F-(aq) + H2O(l)

*Note that this does not represent REDOX, it is simply here to demonstrate the concept of Full/Net Ionic equations until I or someone else can find/think of a better one.

Net Ionic Equations represent all species that undergo a change during the

reaction, whether it be in charge, phase or if it was/becomes part of a weak electrolyte. As with before, weak electrolytes remain unsplit. Anything that does not undergo a change is omitted*. These are called spectator ions.
Example : $HF(aq) + OH^-(aq) F^-(aq) + H_2O(l)$

*Note that $Na^+(aq)$ is not in the net ionic yet $F^-(aq)$ is despite Fluorine being on both sides in aqueous form. This is because $Na^+(aq)$ appears the same way on both sides, but fluorine is part of a weak acid (And thus, a weak electrolyte) $HF(aq)$ as a reactant, but appears by itself as $F^-(aq)$ on the product side. It is because it does not appear the exact same way twice that it is included.
By looking at the reaction and the Net Ionic it becomes easier to determine what the half reactions will be since it is easier to see what is reduced and what is oxidized. (A single species can be both reduced AND oxidized, usually a reactant.)

ELECTROCHEMISTRY

https://www.youtube.com/watch?v=A0KaYHpPxqg
This is pretty helpful for Electrochem

2 Of Amerikz Most Wanted

Two major 'cell' types
Galvanic, or voltaic, thermodynamically favored
Electrolytic, thermodynamically UNfavored (has to use another power source, such as a battery)
Depends on the ΔG
If ΔG is less than 0 ($G<0$) it is FAVORED
If ΔG is greater than 0 ($G>0$) UNFAVORED

<u>Electrodes</u>
Anode
Electrode that gets oxidized

Seems to get smaller

Cathode
Electrode that get reduced
Seems to get larger

Salt bridge
Used to maintain an electric neutrality in a galvanic cell
Anions from the salt bridge flow towards the anode
Cations from the salt bridge flow towards the cathode
Voltmeter is used to measure the cell potential
Electron always flow from anode to cathode
Mnemonic devices
AN OX (Anions Oxidize)
RED CAT (Reductions are for Cations)
FAT CAT (Electrons flow From the Anode To the CAThode)
EPA
Electrolysis has a Positive Anode since the polarities are reversed!
AN OX chases the RED CAT

Cell potential
Positive cell potential makes the cell work
E^0 = E(reduction) - E(oxidation) (Hess's Law type question: change the sign of oxidation half-reaction because it's reversed in the overall redox reaction)
^{0} in E^0 means standard conditions: 298 K, 1 atm, 1.0M in each solution
E^0 is intensive, so it doesn't change with mass. Multiplying a half-reaction by 2 to balance the redox reaction does NOT mean you multiply the E value.
If both reduction potentials positive: the lower potential will probably become the oxidation reaction
If one positive, one negative: the negative one will probably become the oxidation reaction
If they are both negative, the lower negative one will flip and become the oxidation reaction
E(cell) is E^0 at different temperatures, pressures, and molarities

Nernst Equation: E(cell) = E^0 - RT/nF (ln Q)
E^0 is standard reduction potential
R = 8.314 J / mol K
n= moles of electrons transferred (balance redox to see this)
F = 96485 Coulombs per mol of e-
Q = reaction quotient, molarities of **product over** molarities of **reactants** (see equilibrium)
Nernst Equation at 298 K: E(cell) = E^0 - 0.0592/n ln(Q)
Nernst Equation at Equilibrium: E(cell) = -RTln(k)

At equilibrium, the standard cell stops producing voltage so E^0 becomes 0

Concentration cell: a cell with the same elements on either side, driven by molarity difference
Galvanic Cell Notation: Cu | Cu2+ || Cd2+ | Cd
Where Cu(s) is being oxidized in the anode and Cd(s) is being reduced in the cathode
Electromotive force to pull electrons as they move from anode to cathode
Known as E(cell)
Measured in Volts
1V (volt)= 1 Joule/ 1 coulomb
Electrolysis - dimensional analysis questions.
Amps are C/second
Faraday constant 96485 is C/mol e-
Sec * (C/second) * (mol e-/C) * (mol A/mol e-) * (g A/mol A)
Can be applied backward to go from grams or moles of A to time
Just fencepost it out

ΔG, Gibbs Free Energy (kJ/mol)
$\Delta G=\Delta H-T\Delta S$
ΔH=Enthalpy Change, T= Temperature, ΔS= Entropy Change
Important to remember that ΔH is normally in kJ and ΔS is normally in J, so you need to convert it to J for the equation to work
$\Delta G=-nFE$
F=96,485 C per mole of e- , n=mol of electrons transferred (in balanced full equation, usually harder than you'd expect to find), E=reduction potential (Ecell)

Stoichiometry
When dealing with trying to find out information on either the reactant or product side for questions, follow these steps
Convert to moles (grams/molar mass)
Find the limiting reactant (which material runs out first (not always what is the lowest))
Use the balanced equation to see how many moles it will need
Convert moles into desired unit

ENTHALPY

Enthalpy- measure of how much energy is released or absorbed
Energy is basically heat, so it is good to relate it to heat

Basic Rules of enthalpy

When Bonds are MADE, energy is RELEASED

When Bonds are BROKEN, energy is ABSORBED

It takes energy to break bonds (think that atoms are attracted to each other while in bonds, so you need to push them away from each other), making bonds releases energy

Enthalpy is usually written as ΔH

When ΔH is negative ($-\Delta H$) heat is being RELEASED and it is Exothermic

Heat is treated as product (for example a candle flame)

When ΔH is Positive (ΔH) Heat is being ABSORBED and it is endothermic

Heat treated as reactant (ex: melting an ice cube)

Exothermic- heat (energy, or electrons) is released (hot stuff) (PE of Reactants should be higher than PE of products!)

** More likely to be spontaneous

Endothermic- heat(energy, or electrons) is absorbed (cold stuff) (PE of reactants should be lower than PE of products!)

MEMORIZE THESE

Activation energy- is how much energy is needed to fully break all the bonds in the reactants (none form in the products yet)

The top of the peak is the activated complex, and no bonds are present when at it, immediately after product bonds form, immediately before, reactant bonds break

Catalyst- a substance that speeds up a reaction, while not affecting its outcomes, its more of a short cut for the reaction (lowers it for both forward and reverse)

It speeds up a reaction by providing a lower activation energy without changing the products or delta H.

KINETICS

RATE LAWS

Rate Laws- The generic equation for a rate law is Rate = k[A]^x[B]^y[C]^z, with x being the exponent determined by the ‘order’ of [A], y for [B], and z for [C]. Rate laws give us the rate/‘speed’ of the reaction in relevance to the concentrations of the reactants in the ‘rate determining step.’

Rate determining/Slow step- This the part of the reaction that is responsible for how long it will take overall - aka it determines the rate.

Fast Step(s) - Usually we can just ignore these since they are assumed to happen so quickly that they have a negligible impact on the overall reaction time.

ZERO ORDER: FIRST ORDER: SECOND ORDER:

Zero Order Rates- [A] vs time graph linear

First Order Rates- ln[A] vs time graph linear; <u>half life independent of concentration</u>

Second Order Rates- 1/[A] vs time graph linear

Rate Constant k- increases with temperature; units for k found by substituting the units (M for [A] and M/s for Rate) into the rate law than solving for k.

Half Life : $t_{1/2} = 0.693/k$ (only for first order) (In the reference table!) OR can be determined from a graph if given
Concentration vs Time - even if it is first order! (Just look for where the concentration is half of its original value, and to double check your answer, either plug it into the equation and solve for k, or check from the half value and halve it again to see if it is the same amount of time.)

- The half life of a reaction is the time it takes for the concentration of the reactants/species in question to reach half of its original value. (Ex : In a generic equation A + B C where [A] is 1.00 M, the value for the half life would be how long it will take for [A] to decrease from 1.00 M to 0.50 M, from 0.50 M to 0.25 M, and so on and so forth.)

- Half life remains constant throughout the reaction (Only if First-order)! It should not be changing - if you check your answer at different intervals on a graph, and it has changed by a significant amount, or if the half life equation and graph derived values are very different you are doing it wrong! (Sometimes finding an answer by the graph and finding the answer using k can give slightly different values. The same goes for the reverse, if using half life to solve for k. Don't worry if the answer is only changed by .005 or so.)

Beer-Lambert Law- $A=\varepsilon bc$
A - Absorbance
-
ε- molar absorptivity (usually a constant) - specific to the solute and optimal wavelength
b- path/cuvette length/width
c- concentration

Reaction Mechanisms

Collision theory

A postulate that states that a reaction occurs when the reactant molecules collide with **sufficient energy** and **proper orientation**.

There are 5 ways to increase the speed at which the reaction occurs

1. Temperature increase: Means the average kinetic energy of the molecules is higher so the molecules are more likely to collide with enough energy to overcome the activation energy.
2. Surface area increase
3. Catalysts
4. Increasement of concentration: More reactant molecules present means more collisions which in turn raises the likelihood of having successful collisions
5. Pressure increase: Less space for reactant molecules to move around so they hit each other more frequently, resulting in more successful collisions overall

THERMODYNAMICS

ΔH= Enthalpy (Heat/energy)
ΔS= Entropy Change (how chaotic something is)
ΔG= How much free energy a reaction has to give

Basic rules (@ Standard Conditions= °)
All gases are 1 atm
All liquids are pure
All solids are pure
All solutions are at 1 Molarity
Temperature is usually room temp 25 Celsius, 298 Kelvin
Element in normal state has zero energy

ΔH = Enthalpy
$-\Delta H$ is exothermic, more stable
$+\Delta H$ is endothermic, less stable
Pure element should have 0 ΔH

Entropy (ΔS)
Entropy of a perfect, pure crystal @0K=0
Universe prefers chaos (2nd law of thermodynamics)
Gas is the most chaotic state, aqueous being 2nd most chaotic
Solid is the least chaotic state
Increase in Entropy
Increasing in the number of moles of gas
Ex: $A(g)+B(l)2\ C(g)$
Moving from a more ordered state to a more chaotic (solid→liquid→gas)
Negative ΔS means that the system is becoming more orderly, positive ΔS means that the system

is becoming more disorderly

Hess's Law- overall enthalpy change in a reaction is the sum of all the reactions in the process

$q=mc\Delta T$

Used to find heat of process where
q: thermal energy
m: mass
c: Specific heat→ amount of heat required to raise the temperature of 1g of substance by 1°C
ΔT: change in temperature (Tfinal-Tinitial)

Gibbs Free Energy Equation: $\Delta G= \Delta H- T\Delta S$

Enthalpy in solutions
Bonds between ionic substances require energy, but can usually be done with energy made through dipole attractions
This is explained through 3 steps, NaCl will be an example
Solvent Bonds Break, NaCl breaks into Na^+ and Cl^-,
H's and O's in H_2O, which are normally attracted to each other (H^+'s like to naturally be closer to O^-'s) will spread apart more THEY DO NOT BREAK
Then Na^+ moves in between the molecules, and O^-'s are attracted to it, so they huddle around it, the H^+'s huddle around Cl^-

EQUILIBRIUM
Deals with reversible reactions (as opposed to going to completion)
A system is at equilibrium when the forward rate and the reverse rate are equal (products and reactants continue to be produced)
The equilibrium constant Kc is the ratio of the product to reactant at equilibrium
Q tells the reaction progress
Q>K - reaction will proceed to the reverse direction
Q<K - reaction will proceed in the forward direction
Build timeline
Always shift towards equilibrium
______________K<------------Q________________________________

LE CHATELIER'S PRINCIPLE
Whatever stress added to the system will be relieved by the system

***Changing pressure with inert gas will not shift the reaction
***there is an error in this chart - increasing temperature of system favors endothermic reaction

Equilibrium and pressure

Known as Kp as is very interesting when applied to Le Chatelier's principle
Goes to the side with the fewest molecules, when P is increased**
Volume change is analogous to pressure change but they are inversely related
VERY IMPORTANT EQUATION!!!!
Kp = Kc(RT)Δn
Where R is the gas constant (0.08206 L*atm/mol*K) and deltaN is the difference in moles between reactants and products. deltaN can be negative
This convert pressure to concentration and vice versa

ACIDS AND BASES

Strong acids ionize completely
HCl, HBr, HI, *H2SO4, HClO4, HNO3
*H2SO4 is polyprotic, only the first H ionizes completely, the second H is weak
All other acids are weak
Strong bases include hydroxides of Group 1A and most of Group 2A metals
Weak bases usually contain a nitrogenous base

Kw=1.0*10-14 at 25 ^{O}C
Ka * Kb = Kw
Kw is the water dissociation constant.
Since the ionization of water is an endothermic process, higher temperature increases Kw and decreases the pH
Lower pH of pure water does not mean the the water becomes acidic at higher temperature relates to the equation for the autoionization of water, 2H2O H3O + OH . As temperature is raised, the pH of water decreases since it will begin to favor the products, but it remains neutral because pOH also decreases due to the increasing concentration of both ions.)
Neutrality is when [H3O+] = [OH-] (NOT always 7!)

CONJUGATE ACIDS AND BASES

Bronsted-Lowry
The acid is the proton donor and the base is the proton acceptor
This proton is usually shown as an H+ ion, but since H+ does not exist simply by itself, it piggybacks onto an H2O molecule to form the hydronium ion: H3O+ (H+ is acceptable in AP)
Water is amphoteric (can act as either an acid or a base). Another name is amphiprotic
For weak acids, higher percent dissociation for lower concentrations
%dissociation= ([H+]/[HA]) x 100%
For diprotic and triprotic acids, each subsequent dissociation is less favored (Ka1>Ka2>Ka3)
Can be seen in H2SO4 where the first H+ leaves easily but the Ka for HSO4- is much lower (~10^-2)

Oxyacids- more Oxygens will cause the electrons to be pulled toward oxygens and it is easier for the H+ to ionize (more oxygens=stronger acid)

Lewis Model

Acids will accept an electron pair and bases will donate an electron pair.

Weak acids have a strong conjugate base and low Ka value.

Strong acids have weak conjugate base and have a high Ka value

BUFFER SOLUTIONS

Usually made of a weak acid and its conjugate base or vice versa

To make a buffer solution, choose a conjugate acid-base pair with pKa close to the target pH

A buffer will counteract large changes in pH

Capacity of a buffer is dependent on the concentrations of the 2 components of the buffer (higher the concentration, the higher the capacity)

pH= pKA+ log(Base/Acid) OR pOH=pKb + log [HB+]/[B]

k for the HH equ our teacher taught

[H+]=Ka(mol acid/mol base)

Ideal buffer is equal concentration acid/conj. Base and vice versa (pH=pKa)

TITRATION

If you have a solution with unknown concentration and a solution with KNOWN concentration you can perform a titration. This involves putting the known solution (acid or base) into a buret over a flask of your unknown solution.

For best results you want to use an acid-base indicator with a range close to the equivalence point. This way you will be able to see the dramatic change in pH via the change in the indicator color. The most common indicator is phenolphthalein which is colorless in acid and pink in base.

Titrant: The substance with the known molarity which is used to find the molarity of the other substance, usually added to the analyte (our teacher did show us a released problem where analyte/substance w/unknown molarity was being added to the titrant)

Analyte: Substance which is being analyzed to find its molarity

Indicator: pH=pKa

TITRATION GRAPH

Equivalence point is where moles of both the acid and the base are equal in 1:1 ratio

Strong/strong combinations will always have an equivalence point of pH=7.

Weak/strong combinations will have equivalence points with pH higher or lower than 7, leaning towards the stronger acid/base

In a weak acid vs strong base titration, at the half-equivalence point [HA]=[A-] and pH=pKa (also known as a perfect buffer)

Also , if pKa<pH, then the deprotonated form will dominate…a question from last year’s frq that many tripped on

(If the **pH** is lower than the **pKa**, then the compound will be **protonated**. If the **pH** is higher than the **pKa**, then the compound will be **deprotonated**)
$pKa>pH$ = protonated
$pKa<pH$ = deprotonated

pH of an acidic/basic salt after titration (at eq point)
Only strong acid and weak base (or vice versa) will form non-neutral salt
moles acid=moles base=moles of salt produced
If given Kb, find Ka and vice versa
Use RICE table to find concentration of salt at eq point
Use your common equilibrium formula to determine the pH

USEFUL RESOURCES
Course Exam and Description, with useful questions and information
2013 Updated AP Chemistry Compilation Exams
2013 Updated AP Chemistry Practice Exam
2014 AP Chemistry Practice Exam
2014 AP Chemistry Free Response, (Answers here)
2015 AP Chemistry Free Response. (Answers here)
2017 IPE Chemistry AP Exam

Important things to sneak into your calculator (which i will be doing):

nernst
vesper
$Kp=Kc(RT)^{\Delta n}$
order reactions
$E = E^\circ - (0.0592/n)\log Q$
arrhenius equation
Raoult's law
$MM=D(RT)/P$
increasing temp by 10 the rate doubles√
[NMI HDL] nevins mom is hella dumb lol (inside joke but it is for bonds strength)
    a mnemonic device to remember the order of strength of bonds and IMFs -
    network covalent, metallic, ionic, hydrogen bonding, dipole-dipole, london dispersion
Keeping these may help you but they may have no effect. Better safe than sorry!

    (6.022×10^{23})

Convert Moles to Grams (vice versa)
    Moles = (gram/molar mass)
    Grams = (Moles x Molar Mass)

Moles of a gas are found with n=PV/RT (Rearranged **Ideal** Gas Law.)
Most likely will occur at STP (Standard Temperature & Pressure)
Pressure = 1 atm
Temperature = 273 K (0 C)
At STP one mole of an *ideal* gas will be at 22.4 Liters.

P = Pressure in atm
V = Volume in litres
n = # of moles
R = Gas Constant (.0821 $\frac{L*atm}{mol*K}$)
T = Temperature in Kelvin (Take Celsius and add 273 to get Kelvin)(Test won't give Fahrenheit)

Moles = (molarity)(liters of solute)

Percent Composition- percent of mass of each element in a compound
Divide each element's molar mass by compounds molar mass
$Ca(NO_3)_2$ = 164.10
Calcium = 1 = 40.08 x 1 = 40.08
Nitrogen = 2 = 14.01 x 2 = 28.02
Oxygen = 6 = 16 x 6 = 96.00
Ca = 40.08/164.10 = 24.42%
N = 28.02/164.10 = 17.07%
O = 96/164.10 = 58.50%d

Empirical Formula

Empirical formula can be found by calculating the moles of each element present, and then dividing all mole values by the smallest number found. The values given by the division will be the empirical formula subscripts.

Electron Configurations

Nucleus is positive, pulling in electrons which are negative, closer they are stronger the attraction is i.e More protons= more attraction, closer electrons= more attraction

Electrons are repelled by other electrons, an electron between a valence and nucleus causes the valence to be weaker (called shielding)

Completed shells are stable,atoms strive for completed shells

Aufbau principle- when building electron configurations, electrons go into shells in order of increasing energy

Pauli Exclusion Principle- two electrons in a shell will not spin the same way

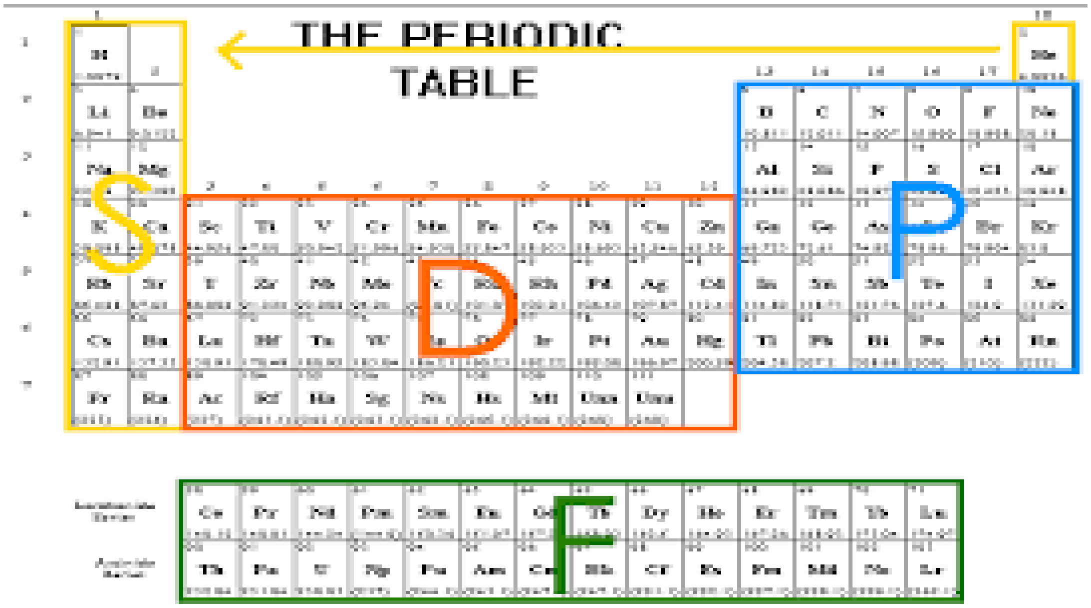
Hund’s rule- when an electron is added it will take up the lowest energy orbital

Amount of energy from an electron is dependent on how far away it is from the nucleus

Bohr Model - Quantum model for the atom where electrons are arranged in specific orbits around the atom - used this to predict the emission spectrum of hydrogen given the electron distribution. Problem - good for predicting some emission spectrums but does not work with most. Also, electrons do not “orbit” in a perfect, flat plane.

If an atom is exposed to electromagnetic levels greater than the ones exerted by the electrons it holds the electrons can be ejected.

Electron Configuration w/ periodic table



Use Periods (horizontal) as first number, then letter of the group, then amount

Examples

Hydrogen: $1s^1$

Helium: $1s^2$

Lithium: 1s^2, 2s^1
Sodium: 1s^2, 2s^2 2p^6, 3s^1
You can also ‘shorthand’ by using the Noble gases as starting points
Argon: [Ne] 3s^2, 3p^6

Moving left to right from periodic table atomic radius (distance from nucleus to valence electrons) decreases (Z_{eff} increases while valence electrons are added to same energy level, so overall the electrons are pulled in more)

Moving down a group, atomic radius increases (more shells mean the valence electrons are further out)

Removing electron creates a positively charged cation (X^{1+}), the energy to remove the first electron **from a gaseous ion** is the ionization energy, the energy to remove the next electron is the second ionization energy, which is always higher than the first ionization energy as the number of protons remains the same while there is one less electron, so less force repelling and the same force attracting, overall requiring more force to remove the electron.

If the electron being removed drops a shell, ionization energy increases MASSIVELY.

Electronegativity
Small atoms have larger electronegativity levels
Atoms close to having stable electron amounts have stronger electronegativity levels

BONDS

Bonds - sharing of electrons, more electrons shared = stronger bond
So single bonds (1 σ) have the least bond energy
Bond Energy Strength: triple (2 π and 1 σ)>double (1 π and 1 σ)>single (1 σ)

Ionic Bond- Bonds held by electrostatic forces, one will be + (cation(red)) while the other is - (anion(black))
Ionic bonds are where the anion completely removes the electron from the cation

Metallic Bond- Metals in the bond are usually stationary and share a sea of electrons, making them good conductors as they give up electrons easily,this also makes the metals malleable

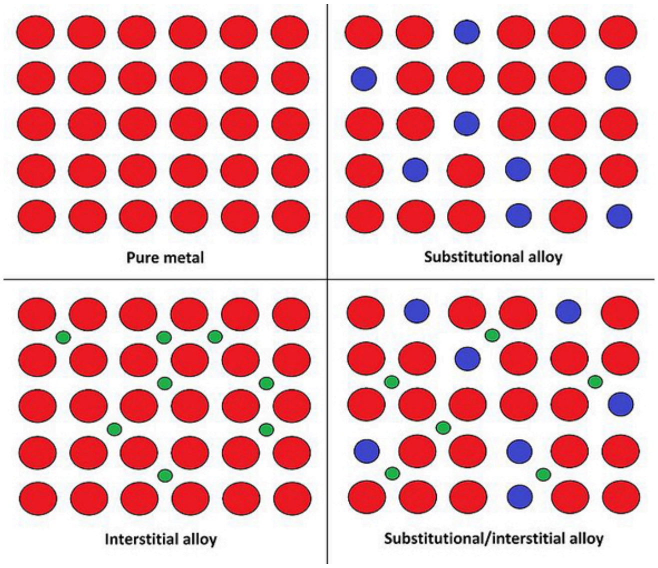
Metals can bond with each other and form alloys

Interstitial Alloy- metals with VERY different atomic radii combine

Example is carbon filling in the cracks around iron to make steel

Substitutional Alloy- Two metals with similar atomic radii combine

Example is Zinc and Copper, it is much more structured and uniform



Covalent Bonds- Two atoms share electrons and both count the electrons as part of their valence shell (nonmetals)

The first covalent bond is called the sigma bond (σ), and has the lowest bond energy and longest bond length, while the following bonds (2nd and 3rd) are called pi bonds

Network Covalent Bonds-these bonds are formed when a group of atoms share covalent bonds into one big electron sharing cluster, it is very hard to break and has a high melting point

Good contenders are Silicon and Carbon since they only have 4 valence electrons (ex. SiO_4)

Polarity- Atoms usually will have a electronic imbalance of sorts due to uneven distribution of electrons (very based on symmetry)

Dipole- The most electronegative element will pull on electrons but not enough to become ionic, but will keep them on one side of the atom, H_2O for instance has 2 H^+ on one side, connected to O^- this makes the O^- side negative, and the H^+ side positive, more charges means more dipole moment, which is how a dipole is measured

Dipole moments can attract each other, causing the H^+'s in H_2O to be attracted (they do not form a bond) to the O-'s in other water molecules (this is a hydrogen bond described below)

Intermolecular forces- forces that exist between 2 or more covalent molecules that hold them together. (ex. London Dispersion forces, dipole/hydrogen bonds.) Stronger IMFs = higher boiling point, lower vapor pressure

Hydrogen Bonds- Special type of dipole moment attraction that involves a positively charged hydrogen to bond to a negative end of another molecule, usually one with high electronegativity (H−F, H−O, H−N).
Extremely strong bond as hydrogen gives up its only electron so it uses all its strength to hold onto it, molecules with these bonds have high boiling/melting points. A group of H-F molecules, or any molecule w hydrogen bonding,are oriented so that the very positive H of one molecule is attracted to the very electronegative F of another molecule.

London Dispersion forces- With large amounts of electrons moving in random patterns it is possible for a majority to move to a specific side, creating a quick bit of polarity (Dipole Moment), which can in turn disrupt other molecules. More electrons means more chance of them congregating (more polarizable), so molecules with high boiling points are the most vulnerable. Larger molecules are more polarizable so have more electrons and will more likely have London dispersion forces
To compare the forces, compare the IMFs - larger is stronger

Bond order = (# of sigma and pi bonds/ # of atoms)
Formal Charge = Group Number - Lone electrons - # of sigma and pi bonds
(Group Number - dots - sticks = formal charge)

- most stable structures generally have **negative** formal charges on the **more electronegative** atoms and **positive** formal charges on the **less electronegative** atoms

Bond Strength + Intermolecular Forces (IMFs)
Strongest to weakest

1. Covalent Bonds
    a. (Usually liquids.)
    b. Network Covalent Bonds are strongest possible (think diamond).
    c. (sharing valence electrons means you need to break 2 sets of valence electrons, making it the toughest to break)
    d. (Melting/boiling point is <u>lower</u> than ionic)
    e. Use 4,5,6 bonds below to base on other covalent bonds strength
    f. Electrons are shared.
    g. Nonmetal-Nonmetal.

2. Ionic Bonds
    a. Usually solid.
    b. Energy to break based on coulombic attraction

    c. Bad conductors in solid form as electrons don't move much.
    d. Generally good conductors in aqueous solution.
    e. Electrons are transferred.
    f. Metal-Nonmetal.
3. Metallic Bonds
    a. Transition metals seem to be strongest of all metals
    b. High melting point, usually below ionic
    c. Generally good conductors as a solid. ('Mobile sea of electrons.')
    d. Metal only.
4. Hydrogen Bonds (Actually an IMF despite 'bond' being in the name.)
    a. Special/specific kind of Dipole attraction
    b. Most commonly between H and F,O,or N ("Hydrogen bonding is FON!")- corresponding F,O, or N must have a pair of unpaired e-
    c. Strongest IMF.
    d. Contains both LDFs *and* DD.
5. Non-Hydrogen Dipole bonds (Dipole-Dipole - DD)
    a. Usually long lasting
    b. Stronger than LDFs but weaker than HB.
    c. Contains LDFs
    d. Polar
6. London Dispersion Forces (Sometimes called Induced Dipole)
    a. Only creates a dipole temporarily
    b. Weakest IMF.
    c. Is the reason why larger molecules tend to have higher boiling points/lower vapour pressure than smaller ones, can be used to help determine which molecules have the stronger IMFs if they both have D-D and/or HB.
    d. Greater # of electrons corresponds to greater polarizability, which makes for stronger LDFs
    e. Every molecule has LDF forces regardless of if they have other IMFs, be it dipole dipole or H bonding.

a

Molecules with weak IMFs (4,5,6 above.) tend to be gases at room temp (Generally have ONLY LDFs)

Molecules with stronger intermolecular forces tends to be liquid at room temp. (Strong LDFs, DD, HB.)

Ionic substances do not have Intermolecular forces. (Usually solid at room temp)

If a molecule in a liquid builds up enough energy it can break free into a vapor, this is vaporization. (This is also responsible for vapor pressure.)

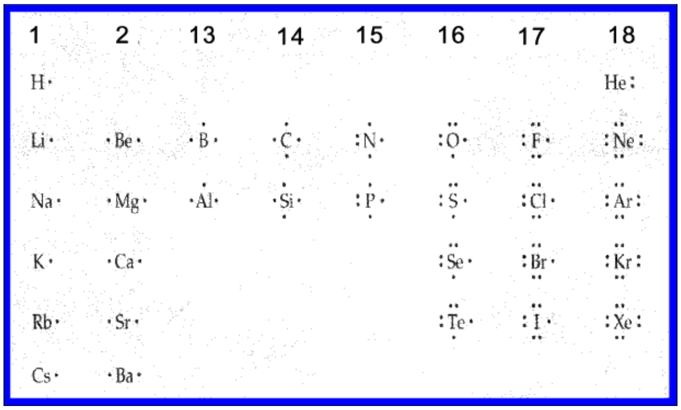
Single < Double < Triple bonds in bond strength

Single > Double > Triple in bond length

Lewis Structures + VSEPR:

It is important to remember that lewis structures are a visual representation of bonds and electrons.

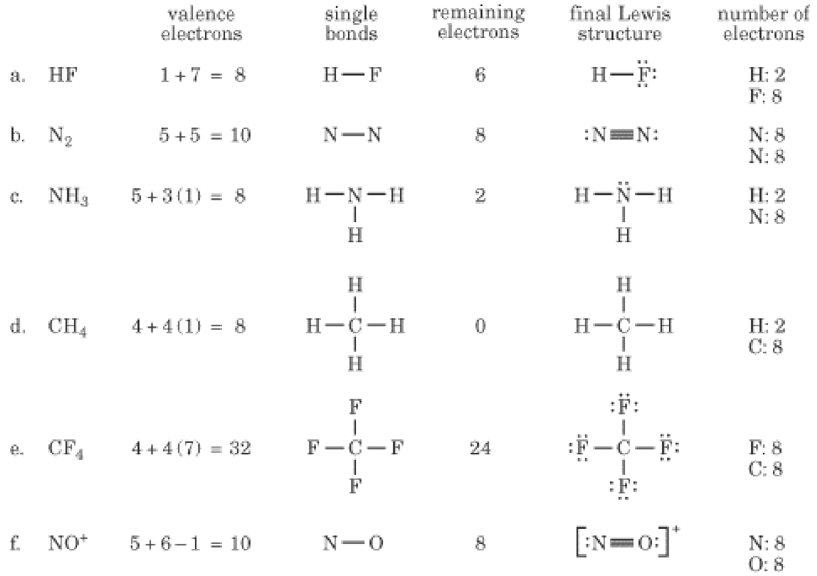
For single atoms/ions all valence electrons should be drawn in. It is important to note that if something is ionized and has no electrons left, then no dots are drawn in - do not go down to the next shell! Also, ions should be put inside of brackets with the charge indicated in the upper right corner.



In molecules, bonds are represented by drawing a line between the two atoms. These lines can be used to represent single, double or triple bonds between two atoms in a molecule.

An important thing to remember is that each bond represents two electrons, any unbonded electrons should be shown in pairs (two dots) (with the exception of an odd total for valence electrons) and that the total number of electrons should be equal to that of the sum of the valence electrons of the atoms. (Make sure to add/subtract the charge from this total if working with polyatomic ions!) (Negative adds electrons, positive subtracts electrons!)

Lastly, as with single atoms, molecules should be drawn inside of brackets if they are ions.



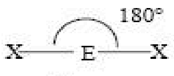
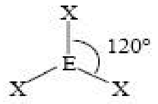
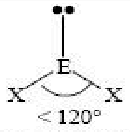
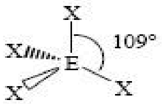
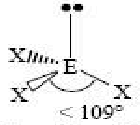
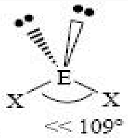
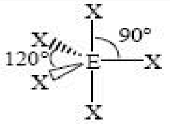
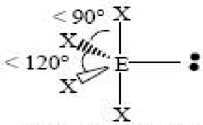
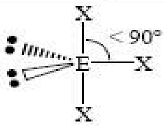
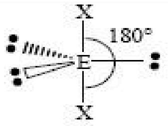
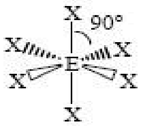
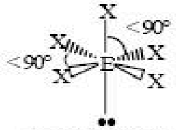
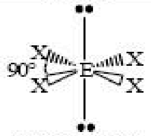
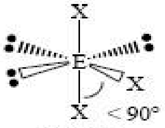
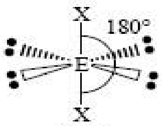
		valence electrons	single bonds	remaining electrons	final Lewis structure	number of electrons
a.	HF	1 + 7 = 8	H—F	6	$H-\ddot{\underset{\cdot\cdot}{F}}:$	H: 2 F: 8
b.	N_2	5 + 5 = 10	N—N	8	$:N\equiv N:$	N: 8 N: 8
c.	NH_3	5 + 3(1) = 8	H—N—H | H	2	$H-\ddot{N}-H$ | H	H: 2 N: 8
d.	CH_4	4 + 4(1) = 8	H | H—C—H | H	0	H | H—C—H | H	H: 2 C: 8
e.	CF_4	4 + 4(7) = 32	F | F—C—F | F	24	$:\ddot{F}:$ | $:\ddot{\underset{\cdot\cdot}{F}}-C-\ddot{\underset{\cdot\cdot}{F}}:$ | $:\underset{\cdot\cdot}{F}:$	F: 8 C: 8
f.	NO^+	5 + 6 − 1 = 10	N—O	8	$[:N\equiv O:]^+$	N: 8 O: 8

VSEPR (Valence Shell Electron Pair Repulsion) Theory : A model used to predict the geometry (structure) of molecules based on bonds and lone pairs of electrons. The idea is that electron pairs will arrange themselves in a way so as to minimise the repulsion between them.

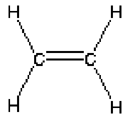
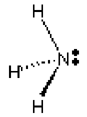
Electron Domain Geometry : The shape of the molecule that arises from the configuration of the molecule about the central atom. The geometry is determined by the number of electron domains (Also called Steric Number.) regardless of whether they are bonds or lone pairs of electrons. **Often referred to as hybridization (only sigma bonds hybridize).**

Molecular Geometry : A more specific shape of the molecule that arises from the configuration of the molecule about the central atom. The geometry is determined by the number of bonds as opposed to the total number of electron domains.

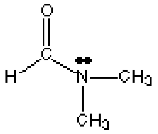
*****BOND ANGLES ARE DETERMINED BY THE ELECTRON DOMAIN GEOMETRY!*****

VSEPR Geometries					
Steric No.	Basic Geometry 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 **Linear**				
3	 **Trigonal Planar**	 **Bent or Angular**			
4	 **Tetrahedral**	 **Trigonal Pyramid**	 **Bent or Angular**		
5	 **Trigonal Bipyramid**	 **Sawhorse or Seesaw**	 **T-shape**	 **Linear**	
6	 **Octahedral**	 **Square Pyramid**	 **Square Planar**	 **T-shape**	 **Linear**

Hybridization Chart

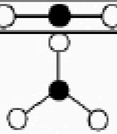
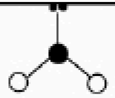
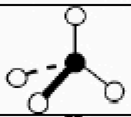
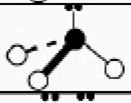
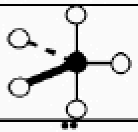
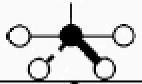
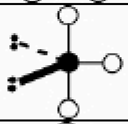
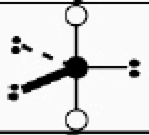
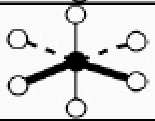
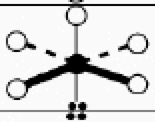
Hybridization	# of "Groups" around hybridized atom	Bond Angle	Geometry	Examples
sp	2	180°	linear	$O=C=O$ $:N\equiv N:$
sp^2	3	120°	trigonal planar	 BF_3
sp^3	4	109.5°	tetrahedral	CH_4  H_2O

Special Case

sp^2	4 - One group is a lone pair and the atom is bonded to a double bond (lone pair always in p orbital)	

Group = sigma bond or unshared pair of electrons

VSEPR SUMMARY

Number of electron pairs around central atom		Full description of the molecule					
BONDING (B)	LONE (E)	Example	Bond angles °	Geometry of Electron Pairs	Geometry of Atoms	3D Shape	Type
2	0	$BeCl_2$	180	Linear	**Linear**		AB_2
3	0	BF_3	120	Trigonal planar	**Trigonal Planar**		AB_3
2	1	SO_2	Slightly less than 120	Trigonal planar	**Bent or V Shaped**		AB_2E
4	0	CH_4	109.5	Tetrahedral	**Tetrahedral**		AB_4
3	1	NH_3	107.5	Tetrahedral	**Trigonal Pyramidal**		AB_3E
2	2	H_2O	104.5	Tetrahedral	**Bent or V Shaped**		AB_2E_2
5	0	PCl_5	120 in plane. 90 perpendicular to plane	Trigonal bipyramidal	**Trigonal Bipyramidal**		AB_5
4	1	SF_4	Complex	Trigonal bipyramid	**Seesaw**		AB_4E
3	2	ClF_3	Approx. 90	Trigonal bipyramidal	**T-Shaped**		AB_3E_2
2	3	XeF_2	180	Trigonal bipyramid	**Linear**		AB_2E_3
6	0	SF_6	90	Octahedral	**Octahedral**		AB_6
5	1	BrF_5	Approx. 90	Octahedral	**Square Pyramidal**		AB_5E
4	2	XeF_4	90	Octahedral	**Square Planar**		AB_4E_2

REACTIONS

Assume molecules without (solid, liquid, or gas) are aqueous (dissolve in water)

Synthesis- When two simple compounds form to make a more complex one

$2H_2 + O_2 \rightarrow 2H_2O$ (liquid)

Decomposition- opposite of Synthesis, complex compound breaking into simple compounds, usually requires energy (heat), usually endothermic

$2MgO \rightarrow 2Mg + O_2$

Acid-Base- when an acid(H+) reacts with a base(OH-) to form a salt and water

$HCl + NaOH \rightarrow NaCl + H_2O$(liquid)

Oxidation-Reduction- reaction that changes the phase of a species (Covered in greater detail later on) (involves transfer of electrons, used in batteries)

$Cu^+ + e^- \rightarrow Cu$(solid)

Precipitation - Two aqueous solutions mix, sometimes a precipitate (solid excess compound) is formed, also a cation and anion mixing can create a salt precipitate (non aqueous)

$K_2CO_3 + Mg(NO_3)_2 \rightarrow 2(KNO_3) + MgCO_3$(solid)

Also since the K+ and NO_3- both start out aqueous and end aqueous, they don't really do much and are called spectator ions

Net ionic equation - break up the molecules (ex. K_2CO_3 into $2K^+ + CO_3^{2-}$), then remove all spectator ions from the equation

The AP test will provide you with most solubility rules, the ones it will not tell you are:

- Compounds with (positive) Alkali metals are always soluble (Li+, Na+,K+ , Rb+, Cs+, Fr+)
- Compounds with NO_3^- are always soluble.
- Compounds with NH_4 are always soluble.
- Chloride (As well as Bromide and Iodide) compounds are usually soluble, the exceptions being when paired with Ag^+, Hg^+, and Pb^+.
- Compounds with SO_4^{-2} are usually soluble, the exceptions being when paired with Ba^{+2}, Sr^{+2}, Hg^{+2}, and Pb^{+2}
- Hydroxides are insoluble except when in a compound with any of the first three bullets.
- CAANGO (chlorates, acetates, ammonium, nitrates, group one)

BALANCING CHEMICAL EQUATIONS

Reactant side (starting compounds go here)→ Product side (compounds after reaction go here)

When balancing an equation, all elements or compounds must be equal (Conservation of mass)

For example, $2(H_2O) \rightarrow 10(H_2) + 5(O_2)$ is incorrect, because the reactant side only has 4 hydrogens, and 2 oxygens

However $2(H_2O) \rightarrow 2(H_2) + O_2$

REDOX

- For a reaction to be a **REDOX** species in the reaction must undergo a **change in charge.**
- Reduction is when electrons are gained.
  - Reduction is Gain: RIG
- Oxidation is when electrons are lost.
  - Oxidation is Loss: OIL
- Just remember OIL RIG
- If an atom is by itself on one side of the reaction, and then on the other side of the reaction it is in a compound that should be an instant clue to you that this is a REDOX reaction, because that means that the charge changes
- Some Important rules for determining charge to follow are:
  - An element by itself always has a charge of 0 (Even diatomics!)
  - Oxygen will always have a charge of -2 (Unless by itself or if it is in the peroxide molecule O2 -2 in which case each oxygen will have a charge of -1.)
  - Hydrogen will have a charge of +1. (Unless by itself, or in the case that it has a -1 charge - don't worry about remembering this, it won't happen unless they tell you it will.)
  - Group 1 and 2 elements usually have a +1 and +2 charge, respectively.
  - Group 17 elements will usually take on a -1 charge.
  - While group 1, 2 and 17 elements usually only have one oxidation state, transition metals can have several.
  - The overall charge of the compound should be zero unless it is an ion, in which case it should then add up to the indicated charge of the ion. (Unless it is a polyatomic you should have memorized, in which case you should have also memorized the charge - it may not be given to you.)
  - Ex :

    Reaction : $AB + C \rightarrow A + CB$

    Let's assume 'A' and 'C' are group one elements and 'B' is a group 17 element. On the reactant side A has a charge of +1, B has a charge of -1, and C is neutral. We know this because A is a group 1, B is a group 17 and C is by itself. On the product side, A now has a charge of 0, B still has a charge of -1, and C now has a charge of +1. We know this because A is by itself, C is a group 1 element, and B is a group 17 element.

This means that:

A has gone from a charge of +1 to 0. **It has been reduced, or has gained electrons/become more negative. (why can't electrons be positive reeeeee)**

B has not undergone a change in charge - **no change means it is not part of the redox!** (This is important in Net Ionic and Half Reactions, defined and explained below.)

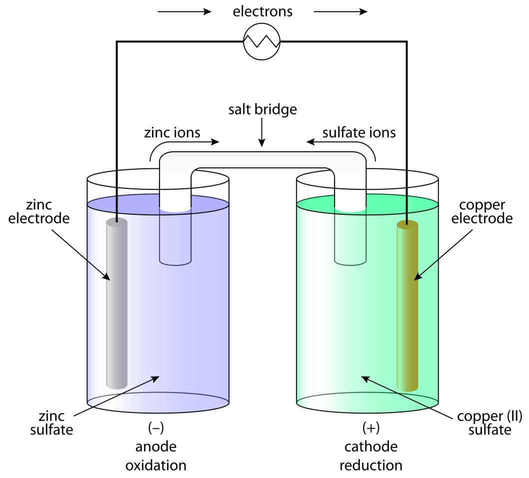
C has gone from a charge of 0 to +1. **It has been oxidized, or has lost electrons/become more positive.**

Balancing REDOX:

- Usually the types of REDOX reactions we should expect to deal with are not so simple as the example given above. **They generally will require balancing the equation as well**, and when this is the case, they usually also need to be broken down into half reactions **so as to maintain both a conservation of charge and mass.**
- Full Ionic Equations represent all species present in the reaction, whether they 'participate' or not. **Weak electrolytes are NOT split up into ions.** (These are also good to use when determining precipitates if you are shaky on solubility rules.)
  - Example :
    $HF(aq) + NaOH(aq) \rightarrow NaF(aq) + H_2O(l)$* will become
    $HF(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + F^-(aq) + H_2O(l)$
    *Note that this does not represent REDOX, it is simply here to demonstrate the concept of Full/Net Ionic equations until I or someone else can find/think of a better one.
- Net Ionic Equations represent all species that undergo a change during the reaction, whether it be in charge, phase or if it was/becomes part of a weak electrolyte. As with before, weak electrolytes remain unsplit. **Anything that does not undergo a change is omitted*. These are called spectator ions.**
  - Example : $HF(aq) + OH^-(aq) \rightarrow F^-(aq) + H_2O(l)$
    *Note that Na+(aq) is not in the net ionic yet F-(aq) is despite Fluorine being on both sides in aqueous form. This is because Na+(aq) appears the same way on both sides, but fluorine is part of a weak acid (And thus, a weak electrolyte) HF(aq) as a reactant, but appears by itself as F-(aq) on the product side. **It is because it does not appear the exact same way twice that it is included.**
- By looking at the reaction and the Net Ionic it becomes easier to determine what the half reactions will be since it is easier to see what is reduced and what is oxidized. (A single species can be both reduced AND oxidized, usually a reactant.)

ELECTROCHEMISTRY

https://www.youtube.com/watch?v=A0KaYHpPxqg
This is pretty helpful for Electrochem



Two major ‘cell’ types

- Galvanic, or voltaic, thermodynamically favoured
- Electrolytic, thermodynamically <u>UNfavoured</u> (has to use another power source, such as a battery)
  - Depends on the ΔG
    - If ΔG is less than 0 (G<0) it is FAVOURED
    - If ΔG is greater than 0 (G>0) UNFAVOURED
- Electrodes
  - Anode
    - Electrode that gets oxidized
      - <u>Seems to get smaller</u>
  - Cathode
    - Electrode that get reduced
      - <u>Seems to get larger</u>
- Salt bridge
  - Used to maintain an <u>electric</u> neutrality in a galvanic cell

    - Anions from the salt bridge flow towards the anode
    - Cations from the salt bridge flow towards the cathode
- Voltmeter is used to measure the cell potential
- Electron always flow from anode to cathode
- Mnemonic devices
    - AN OX (Anions Oxidize)
    - RED CAT (Reductions are for Cations)
    - FAT CAT (Electrons flow From the Anode To the CAThode)
    - EPA
        - Electrolysis has a Positive Anode since the polarities are reversed!
    - AN OX chases the RED CAT

- Cell potential
    - Positive cell potential makes the cell work
    - E^0 = E(reduction) - E(oxidation) (Hess's Law type question: change the sign of oxidation half-reaction because it's reversed in the overall redox reaction)
        - ^{0} in E^0 means standard conditions: 298 K, 1 atm, 1.0M in each solution
        - **E^0 is intensive, so it doesn't change with mass.** Multiplying a half-reaction by 2 to balance the redox reaction does NOT mean you multiply the E value.
        - If both reduction potentials positive: the lower potential will probably become the oxidation reaction
        - If one positive, one negative: the negative one will probably become the oxidation reaction
        - If they are both negative, the lower negative one will flip and become the oxidation reaction
    - E(cell) is E^0 at different temperatures, pressures, and molarities
        - Nernst Equation: $E(cell) = E^0 - RT/nF (\ln Q)$
            - E^0 is standard reduction potential
            - R = 8.314 J / mol K
            - n= moles of electrons transferred (balance redox to see this)
            - F = 96485 Coulombs per mol of e-
            - Q = reaction quotient, molarities of product over molarities of reactants (see equilibrium)
        - Nernst Equation at 298 K: $E(cell) = E^0 - 0.0592/n \ln(Q)$
        - Nernst Equation at Equilibrium: $E(cell) = -RT\ln(k)$
            - At equilibrium, the standard cell stops producing voltage so E^0 becomes 0
    - Concentration cell: a cell with the same elements on either side, driven by molarity difference
    - Galvanic Cell Notation: Cu | Cu2+ || Cd2+ | Cd
        - Where Cu(s) is being oxidized in the anode and Cd(s) is being reduced in the cathode

  - Electromotive force to pull electrons as they move from anode to cathode
    - Known as E(cell)
      Measured in Volts
      1V (volt)= 1 Joule/ 1 coulomb
- Electrolysis - dimensional analysis questions.
  - Amps are C/second
  - Faraday constant 96485 is C/mol e-
  - Sec * (C/second) * (mol e-/C) * (mol A/mol e-) * (g A/mol A)
  - Can be applied backward to go from grams or moles of A to time
  - Just fencepost it out
- ΔG, Gibbs Free Energy (kJ/mol)
  - $\Delta G=\Delta H-T\Delta S$
    - ΔH=Enthalpy Change, T= Temperature, ΔS= Entropy Change
    - **Important to remember that ΔH is normally in kJ and ΔS is normally in J, so you need to convert it to J for the equation to work**
  - $\Delta G=-nFE$
    - F=96,485 C per mole of e- , n=mol of electrons transferred (in balanced full equation, usually harder than you'd expect to find), E=reduction potential (E_{cell})

When dealing with trying to find out information on either the reactant or product side for questions, follow these steps

1. Convert to moles (grams/molar mass)
2. Find the limiting reactant (which material runs out first (not always what is the lowest))
3. Use the balanced equation to see how many moles it will need
4. Convert moles into desired unit

ENTHALPY

Enthalpy- measure of how much energy is released or absorbed
Energy is basically heat, so it is good to relate it to heat

Basic Rules of enthalpy
When Bonds are MADE, energy is RELEASED
When Bonds are BROKEN, energy is ABSORBED
It takes energy to break bonds (think that atoms are attracted to each other while in bonds, so you need to push them away from each other), making bonds releases energy

Enthalpy is usually written as ΔH

When ΔH is negative ($-\Delta H$) heat is being RELEASED and it is Exothermic

- Heat is treated as product (for example a candle flame)

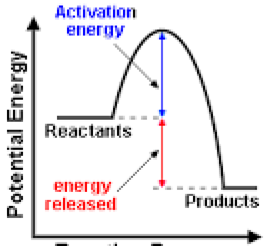
When ΔH is Positive (ΔH) Heat is being ABSORBED and it is endothermic

- Heat treated as reactant (ex: melting an ice cube)

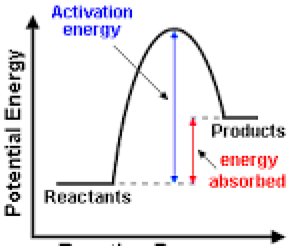
Exothermic- heat (energy, or electrons) is released (hot stuff) (PE of Reactants should be lower than PE of products!)

** More likely to be spontaneous

Endothermic- heat(energy, or electrons) is absorbed (cold stuff) (PE of Products should be lower than PE of reactants!)



Exothermic reaction



Endothermic reaction

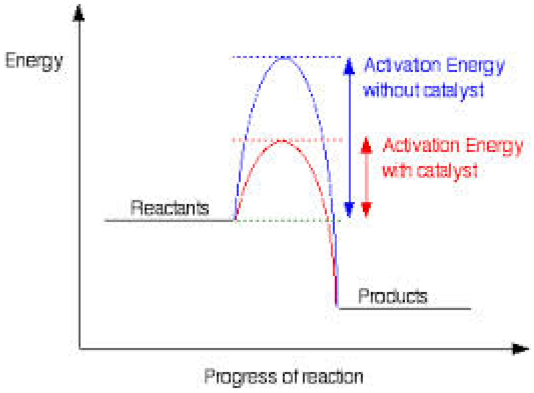
MEMORIZE THESE

Activation energy- is how much energy is needed to fully break all the bonds in the reactants (none form in the products yet)

The top of the peak is the activated complex, and no bonds are present when at it, immediately after product bonds form, immediately before, reactant bonds break

Catalyst- a substance that speeds up a reaction, while not affecting its outcomes, it’s more of a short cut for the reaction (lowers it for both forward and reverse)

<u>It speeds up a reaction by providing a lower activation energy</u> without changing the products or delta H.



<u>KINETICS</u>

RATE LAWS

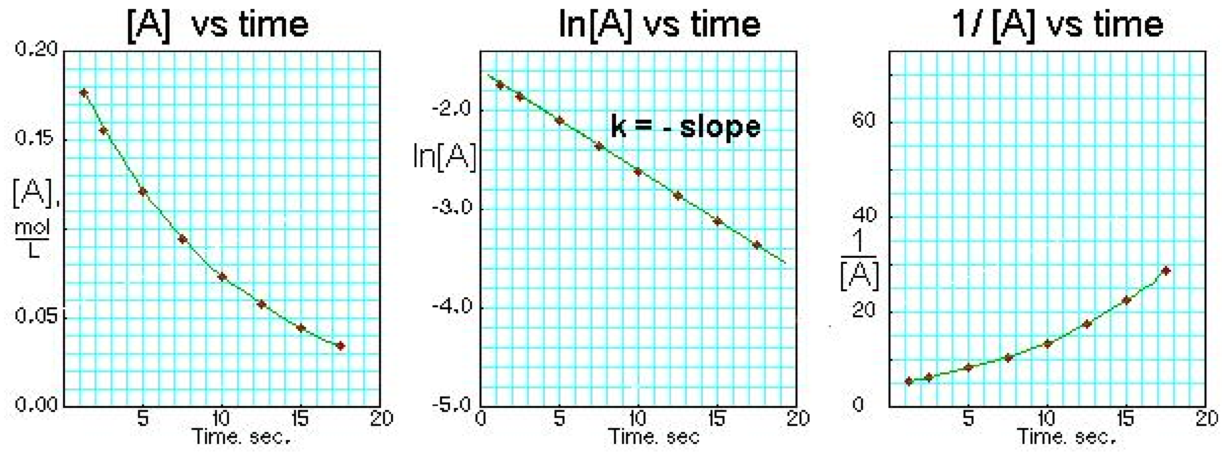
Rate Laws- The generic equation for a rate law is Rate = k[A]^x[B]^y[C]^z, with x being the exponent determined by the 'order' of [A], y for [B], and z for [C]. Rate laws give us the rate/'speed' of the reaction in relevance to the concentrations of the reactants in the 'rate determining step.'

Rate determining/Slow step- This the part of the reaction that is responsible for how long it will take overall - aka it determines the rate.

Fast Step(s) - Usually we can just ignore these since they are assumed to happen so quickly that they have a negligible impact on the overall reaction time.

ZERO ORDER: **FIRST ORDER:** **SECOND ORDER:**

Zero Order Rates- [A] vs time graph linear



First Order Rates- ln[A] vs time graph linear; half life independent of concentration

Second Order Rates- 1/[A] vs time graph linear

Rate Constant k- increases with temperature; units for k found by substituting the units (M for [A] and M/s for Rate) into the rate law than solving for k.

Half Life : $t_{1/2} = \frac{0.693}{k}$ (In the reference table!) OR can be determined from a graph if given Concentration vs Time - even if it is first order! (Just look for where the concentration is half of its original value, and to double check your answer, either plug it into the equation and solve for k, or check from the half value and halve it again to see if it is the same amount of time.)

- The **half life** of a reaction is the **time it takes for the concentration of the reactants/species in question to reach half of its original value.** (Ex : In a generic equation A + B →C where [A] is 1.00 M, the value for the half life would be how long it will take for [A] to decrease from 1.00 M to 0.50 M, from 0.50 M to 0.25 M, and so on and so forth.)
- **Half life *remains constant* throughout the reaction (Only if First-order)!** It should not be changing - if you check your answer at different intervals on a graph, and it has changed by a significant amount, or if the half life equation and graph derived values are very different you are doing it wrong! (Sometimes finding an answer by the graph and finding the answer using k can give slightly different values. The same goes for the reverse, if using half life to solve for k. Don't worry if the answer is only changed by ±.005 or so.)

Beer-Lambert Law- $A=\varepsilon bc$

A - Absorbance

-

ε- molar absorptivity (usually a constant) - specific to the solute and optimal wavelength

b- path/cuvette length/width

c- concentration

Reaction Mechanisms

Collision theory

A postulate that states that a reaction occurs when the reactant molecules collide with sufficient energy and proper orientation.

There are 5 ways to increase the speed at which the reaction occurs

- Temperature increase: Means the average kinetic energy of the molecules is higher so the molecules are more likely to collide with enough energy to overcome the activation energy.
- Surface area increase

- Catalysts
- Increasement of concentration: More reactant molecules present means more collisions which in turn raises the likelihood of having successful collisions
- Pressure increase: Less space for reactant molecules to move around so they hit each other more frequently, resulting in more successful collisions overall

THERMODYNAMICS

ΔH= Enthalpy (Heat/energy)
ΔS= Entropy Change (how chaotic something is)
ΔG= How much free energy a reaction has to give

Basic rules (@ Standard Conditions= °)

- All gases are 1 atm
- All liquids are pure
- All solids are pure
- All solutions are at 1 Molarity
- Temperature is usually room temp 25 Celsius, 298 Kelvin
- Element in normal state has zero energy

ΔH = Enthalpy
$-\Delta H$ is exothermic, more stable
$+\Delta H$ is endothermic, less stable
Pure element should have 0 ΔH

Entropy (ΔS)

- Entropy of a perfect, pure crystal @0K=0
- Universe prefers chaos (2nd law of thermodynamics)
- Gas is the most chaotic state, aqueous being 2nd most chaotic
- Solid is the least chaotic state
- Increase in Entropy
  - Increasing in the number of moles of gas
    - Ex: $A(g)+B(l) \rightarrow 2\ C(g)$
  - Moving from a more ordered state to a more chaotic (solid→liquid→gas)

- Negative ΔS means that the system is becoming more orderly, positive ΔS means that the system is becoming more disorderly

Hess's Law- overall enthalpy change in a reaction is the sum of all the reactions in the process

$q=mc\Delta T$
- Used to find heat of process where
    - q: thermal energy
    - m: mass
    - c: Specific heat→ amount of heat required to raise the temperature of 1g of substance by 1°C
    - ΔT: change in temperature (T_{final}-$T_{initial}$)

Gibbs Free Energy Equation: $\Delta G= \Delta H- T\Delta S$

	$\Delta H < 0$	$\Delta H > 0$
$\Delta S > 0$	Spontaneous at all T ($\Delta G < 0$)	Spontaneous at high T (when $T\Delta S$ is large)
$\Delta S < 0$	Spontaneous at low T (when $T\Delta S$ is small)	Non-spontaneous at all T ($\Delta G > 0$)

Enthalpy in solutions

Bonds between ionic substances require energy, but can usually be done with energy made through dipole attractions

This is explained through 3 steps, NaCl will be an example

1. Solvent Bonds Break, NaCl breaks into Na+ and Cl-,
2. H's and O's in H_2O, which are normally attracted to each other (H+'s like to naturally be closer to O-'s) will spread apart more THEY DO NOT BREAK
3. Then Na+ moves in between the molecules, and O-'s are attracted to it, so they huddle around it, the H+'s huddle around Cl-

EQUILIBRIUM

Deals with reversible reactions (as opposed to going to completion)
A system is at equilibrium when the forward rate and the reverse rate are equal (products and reactants continue to be produced)

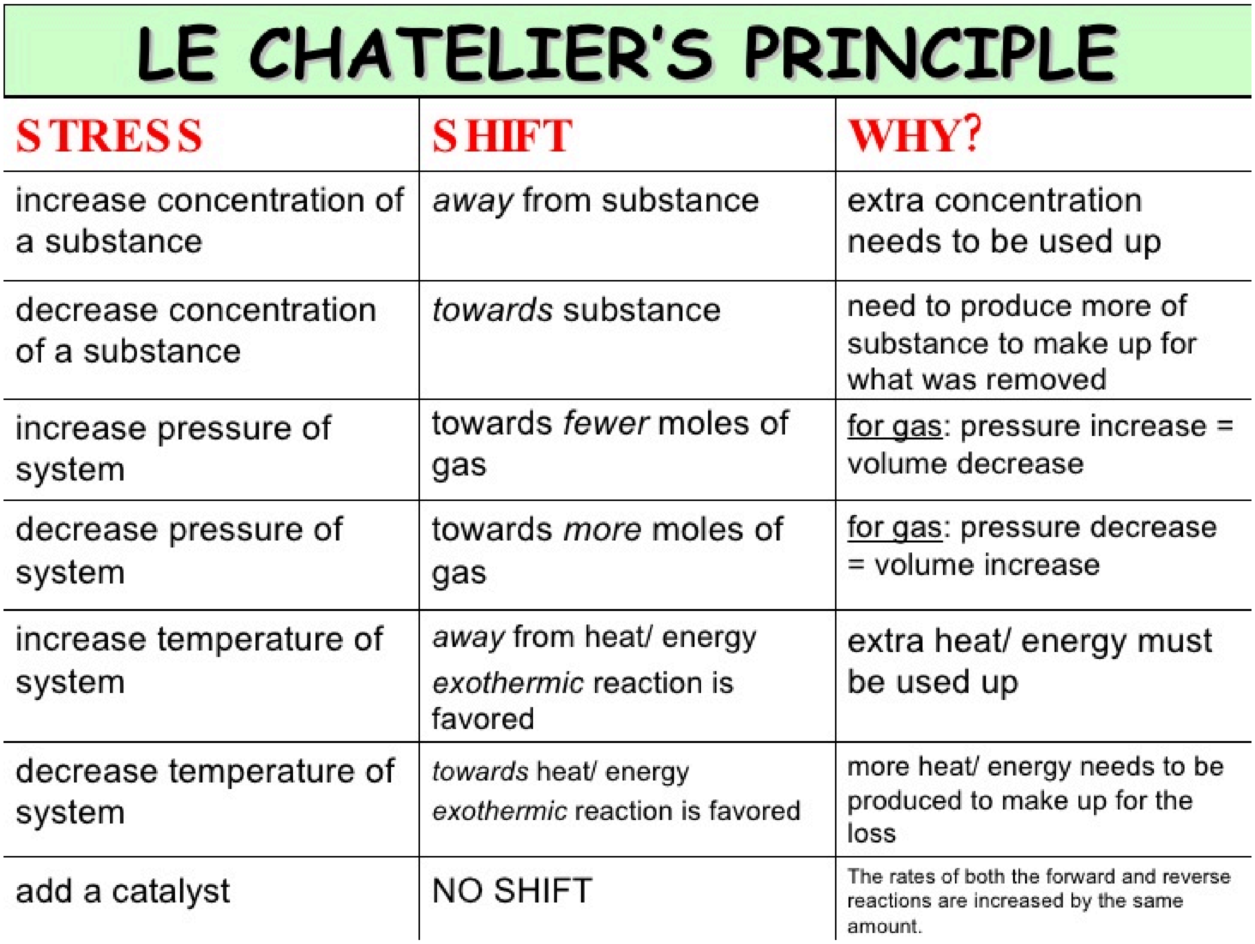
- The equilibrium constant Kc is the ratio of the product to reactant at equilibrium
- Q tells the reaction progress
    - Q>K - reaction will proceed to the reverse direction
    - Q<K - reaction will proceed in the forward direction
    - Build timeline

Always shift towards equilibrium

______________K<------------Q_________________________________

I. LE CHATELIER PRINCIPLE

Whatever stress added to the system will be relieved by the system



LE CHATELIER'S PRINCIPLE

STRESS	SHIFT	WHY?
increase concentration of a substance	*away* from substance	extra concentration needs to be used up
decrease concentration of a substance	*towards* substance	need to produce more of substance to make up for what was removed
increase pressure of system	towards *fewer* moles of gas	for gas: pressure increase = volume decrease
decrease pressure of system	towards *more* moles of gas	for gas: pressure decrease = volume increase
increase temperature of system	*away* from heat/ energy *exothermic* reaction is favored	extra heat/ energy must be used up
decrease temperature of system	*towards* heat/ energy *exothermic* reaction is favored	more heat/ energy needs to be produced to make up for the loss
add a catalyst	NO SHIFT	The rates of both the forward and reverse reactions are increased by the same amount.

***Changing pressure with inert gas will not shift the reaction
***there is an error in this chart - increasing temperature of system favors *endo*thermic reaction

Equilibrium and pressure

- Known as Kp as is very interesting when applied to Le Chatelier's principle
    - Goes to the side with the fewest molecules, when P is increased**
    - Volume change is analogous to pressure change but they are inversely related

VERY IMPORTANT EQUATION!!!!

$$K_p = K_c(RT)^{\Delta n}$$

Where R is the gas constant (0.08206 L*atm/mol*K) and deltaN is the difference in moles between reactants and products. deltaN can be negative
This convert pressure to concentration and vice versa

ACIDS AND BASES

Strong acids ionize completely

- HCl, HBr, HI, *H_2SO_4, $HClO_4$, HNO_3
    - *H2SO4 is polyprotic, only the first H ionizes completely, the second H is weak
    - All other acids are weak

Strong bases include hydroxides of Group 1A and most of Group 2A metals

- Weak bases usually contain a nitrogenous base

K_w=1.0*10^{-14} at 25 ^{O}C

- $K_a * K_b = K_w$
- K_w is the water dissociation constant.

Since the ionization of water is an endothermic process, higher temperature increases K_w and decreases the pH

- Lower pH of pure water does not mean the the water becomes acidic at higher temperature relates to the equation for the autoionization of water, $2H_2O \leftrightarrow H_3O + OH$. As temperature is raised, the pH of water decreases since it will begin to favor the products, but it remains neutral because pOH also decreases due to the increasing concentration of both ions.)
- Neutrality is when $[H_3O^+] = [OH^-]$ (NOT always 7!)

CONJUGATE ACIDS AND BASES

Bronsted-Lowry

- The acid is the proton donor and the base is the proton acceptor
- This proton is usually shown as an H^+ ion, but since H^+ does not exist simply by itself, it piggybacks onto an H_2O molecule to form the hydronium ion: H_3O^+ (H^+ is acceptable in AP)
- Water is amphoteric (can act as either an acid or a base). Another name is amphiprotic
- For weak acids, higher percent dissociation for lower concentrations
- %dissociation= ([H+]/[HA]) x 100%
- For diprotic and triprotic acids, each subsequent dissociation is less favored (Ka1>Ka2>Ka3) Can be seen in H2SO4 where the first H+ leaves easily but the Ka for HSO4- is much lower (~10^-2)
- Oxyacids- more Oxygens will cause the electrons to be pulled toward oxygens and it is easier for the H+ to ionize (more oxygens=stronger acid)

Lewis Model

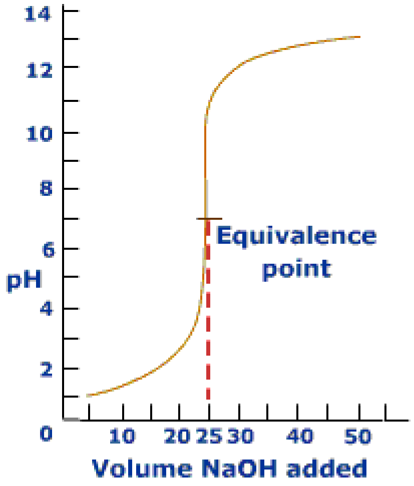
- Acids will accept an electron pair and bases will donate an electron pair.
- Weak acids have a strong conjugate base and low Ka value.
- Strong acids have weak conjugate base and have a high Ka value

BUFFER SOLUTIONS

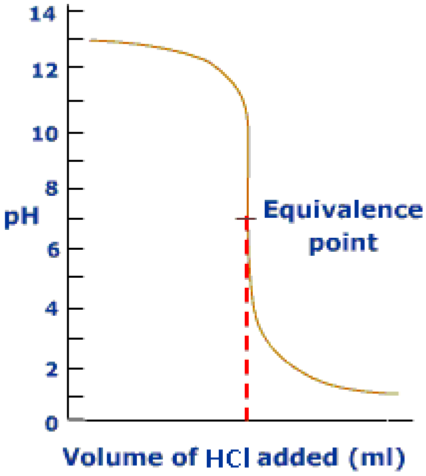
- Usually made of a weak acid and its conjugate base or vice versa
- To make a buffer solution, choose a conjugate acid-base pair with pKa close to the target pH
- A buffer will counteract large changes in pH
- Capacity of a buffer is dependent on the concentrations of the 2 components of the buffer (higher the concentration, the higher the capacity)
- $pH = pK_A + \log(\frac{[Base]}{[Acid]})$ OR $pOH = pKb + \log [HB+]/[B]$
- k for the HH equ our teacher taught
- $[H+] = Ka(\text{mol acid}/\text{mol base})$
- Ideal buffer is equal concentration acid/conj. Base and vice versa (pH=pKa)

TITRATION

- If you have a solution with unknown concentration and a solution with KNOWN concentration you can perform a titration. This involves putting the known solution (acid or base) into a buret over a flask of your unknown solution.
- For best results you want to use an acid-base indicator with a range close to the equivalence point. This way you will be able to see the dramatic change in pH via the change in the indicator colour. The most common indicator is phenolphthalein which is colorless in acid and pink in base.



- Titration curve of strong acid (HCl) with a strong base (NaOH)



Titration curve of strong base (NaOH) with strong acid (HCl)

- Titrant: The substance with the known molarity which is used to find the molarity of the other substance, usually added to the analyte (our teacher did show us a

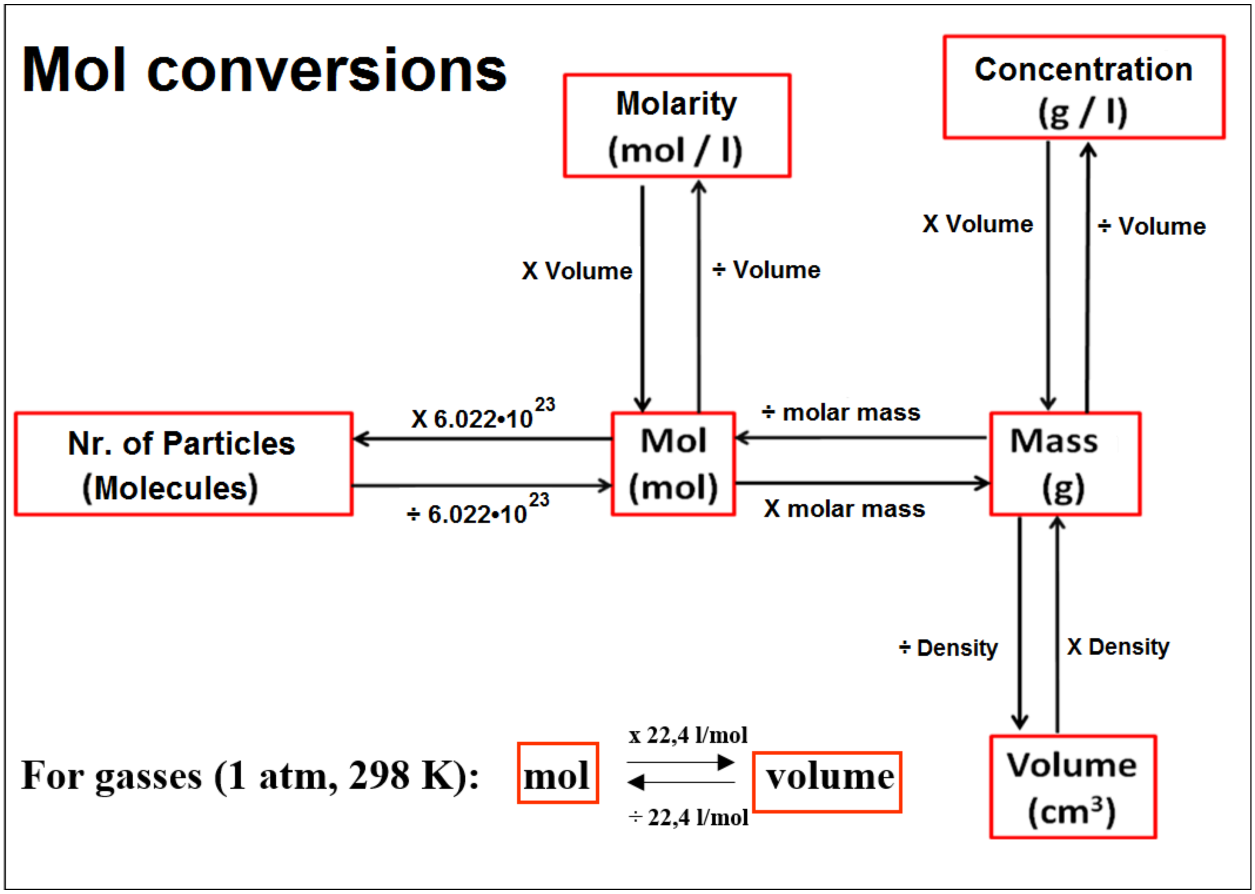
released problem where analyte/substance w/unknown molarity was being added to the titrant)
- Analyte: Substance which is being analyzed to find its molarity
- Indicator: pH=pKa
- Useful Titration Calculator: Used to find the pH of any mixture, the volume or concentration to reach a specific pH/equiv. point/midpoint. Also shows the titration curve graph, and an optional step-by-step solution.

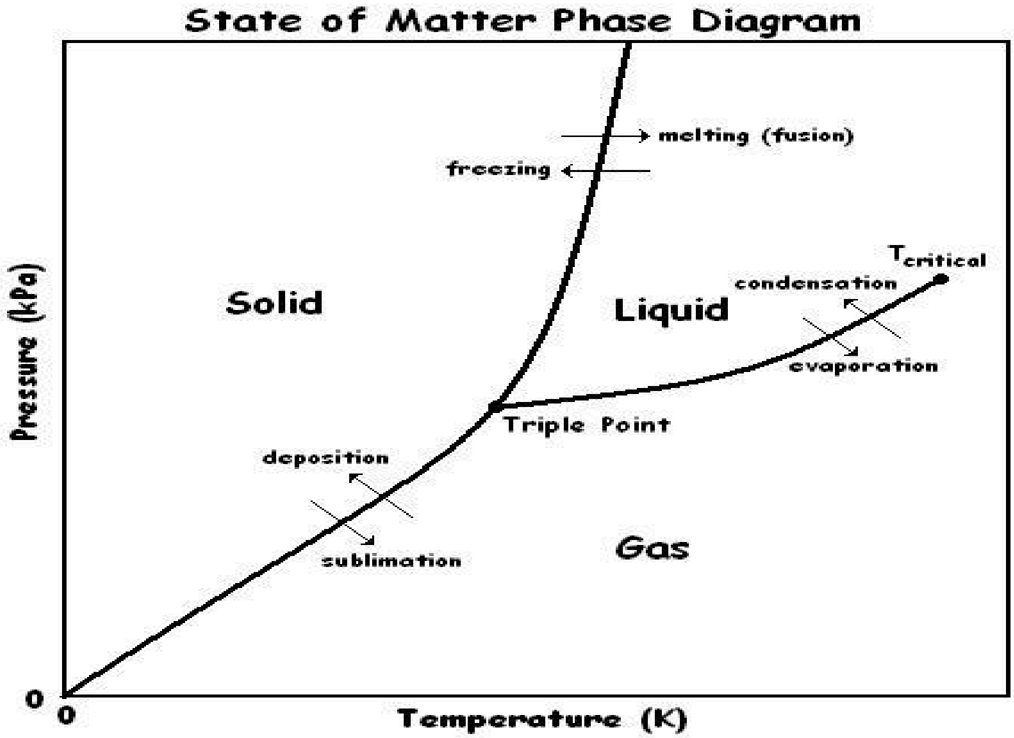
TITRATION GRAPH
- Equivalence point is where moles of both the acid and the base are equal in 1:1 ratio
- Strong/strong combinations will always have an equivalence point of pH=7.
- Weak/strong combinations will have equivalence points with pH higher or lower than 7, leaning towards the stronger acid/base
- In a weak acid vs strong base titration, at the half-equivalence point $[HA]=[A^-]$ and pH=pKa (also known as a perfect buffer)
- Also , if pKa<pH, then the deprotonated form will dominate…a question from last year's frq that many tripped on

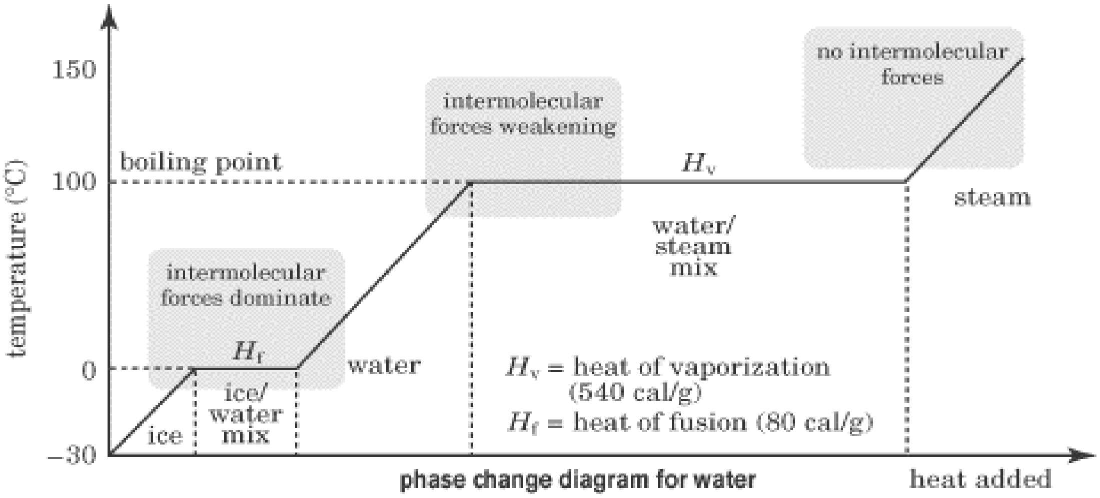
pH of an acidic/basic salt after titration (at eq point)
- Only strong acid and weak base (or vice versa) will form non-neutral salt
- moles acid+moles base=moles of salt produced
- If given K_b, find K_a and vice versa
- Use RICE table to find concentration of salt at eq point
- Use your common equilibrium formula to determine the pH

DELICIOUS AND USEFUL RESOURCES
- Course Exam and Description, with useful questions and information
- 2013 Updated AP Chemistry Compilation Exams
- 2013 Updated AP Chemistry Practice Exam
- 2014 AP Chemistry Practice Exam
- 2014 AP Chemistry Free Response, (Answers here)
- 2015 AP Chemistry Free Response. (Answers here)
- https://drive.google.com/drive/folders/1qNkNQfORqCUPrDDKsxfYKZx3DqbUIr5L
  (This one ^ has all AP IPE tests from 2017, you'll have to find the Chem one)
- https://www.adriandingleschemistrypages.com/ap/2017-ap-chemistry-exam-draft-answers-and-comments/
- Detailed review over every unit
    - ^ includes videos, slides, notes, and quizzes over every topic/unit

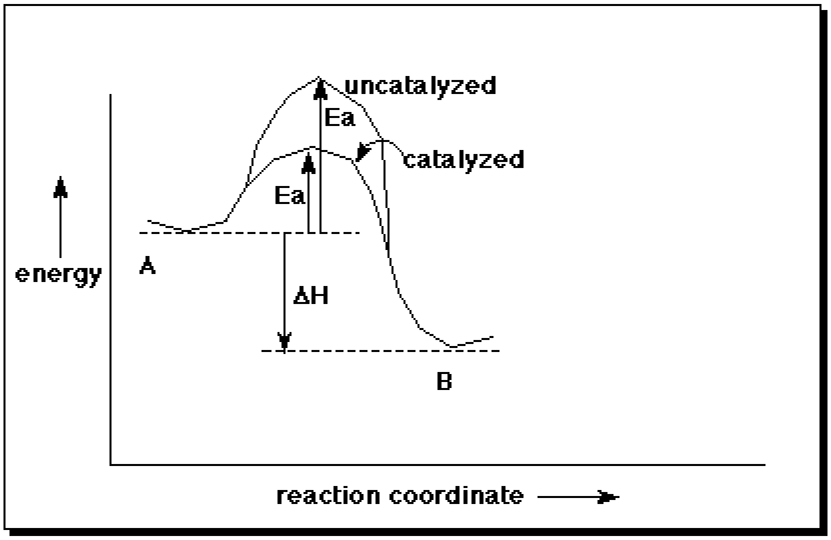
APPENDIX

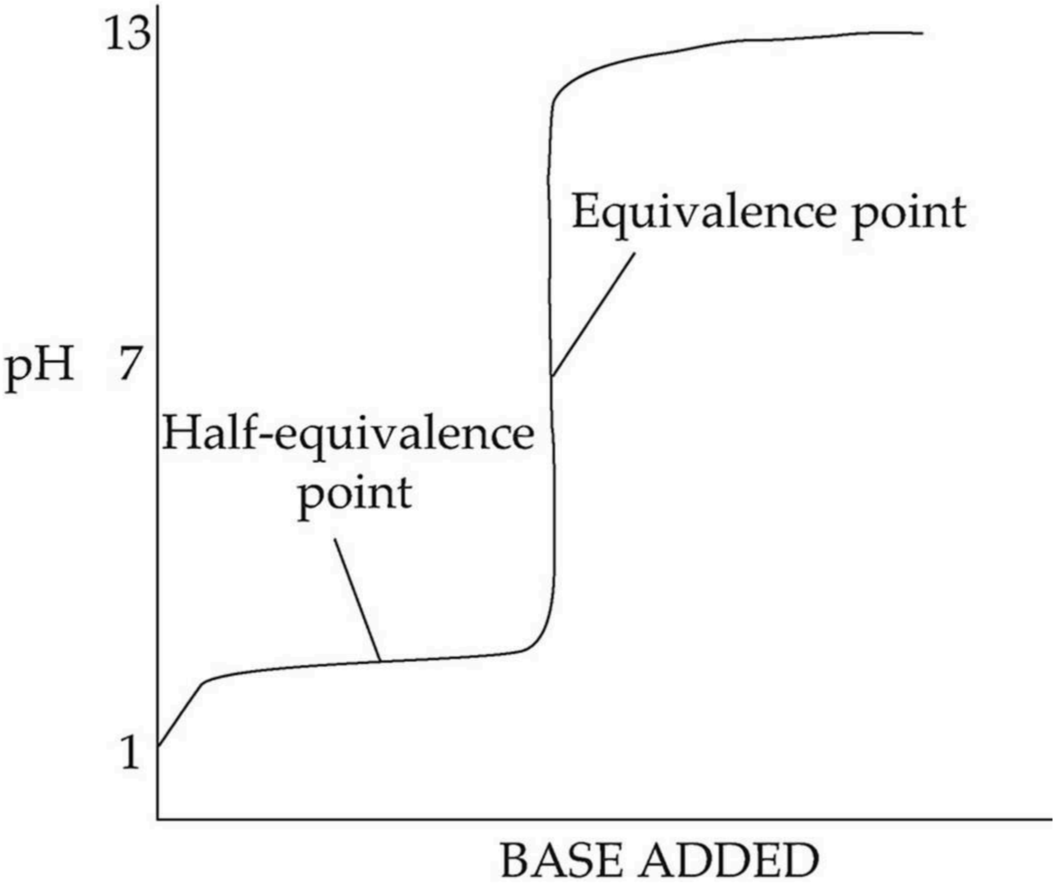

Mol conversions
Molarity (mol / l)
Concentration (g / l)
X Volume
÷ Volume
X Volume
÷ Volume
Nr. of Particles (Molecules)
X 6.022•10^23
÷ 6.022•10^23
Mol (mol)
÷ molar mass
X molar mass
Mass (g)
÷ Density
X Density
Volume (cm³)
For gasses (1 atm, 298 K):
mol
x 22,4 l/mol
÷ 22,4 l/mol
volume

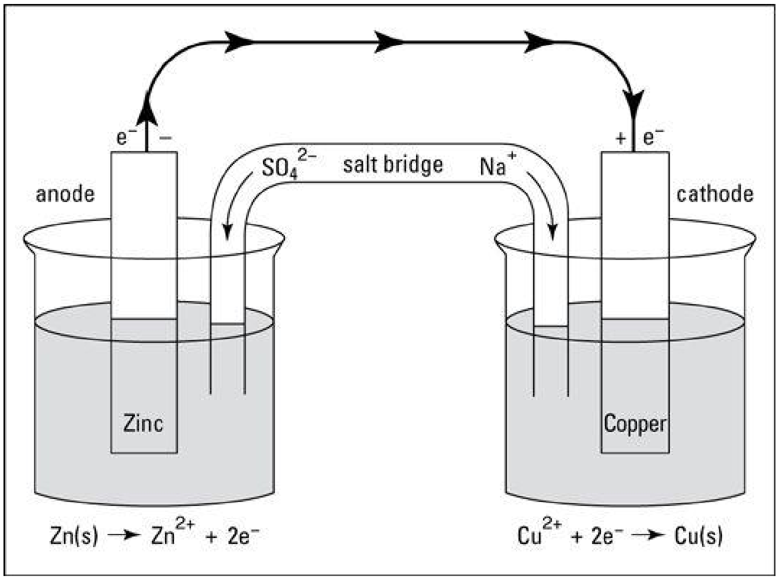

State of Matter Phase Diagram
melting (fusion)
freezing
Solid
Liquid
Gas
T critical
condensation
evaporation
Triple Point
deposition
sublimation
Pressure (kPa)
Temperature (K)
0
0

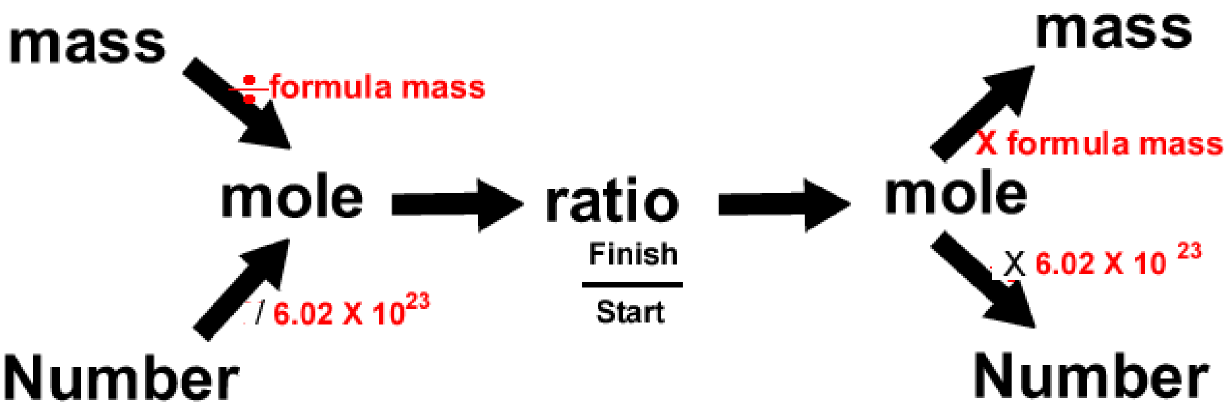


phase change diagram for water


uncatalyzed
Ea
catalyzed
Ea
energy
A
ΔH
B
reaction coordinate


13
Equivalence point
pH 7
Half-equivalence point
1
BASE ADDED





- Important things to sneak into your calculator (which i will be doing):

- nernst
- vesper
- Kp=Kc(RT)^Δn
- order reactions
- E = E° - (0.0592/n)log Q
- arrhenius equation
- Raoult's law
- MM=D(RT)/P
- increasing temp by 10 the rate doubles√
- [NMI HDL] nevins mom is hella dumb lol (inside joke but it is for bonds strength)
  - a mnemonic device to remember the order of strength of bonds and IMFs - network covalent, metallic, ionic, hydrogen bonding, dipole-dipole, london dispersion

Keeping these may help you but they may have no effect. Better safe than sorry!