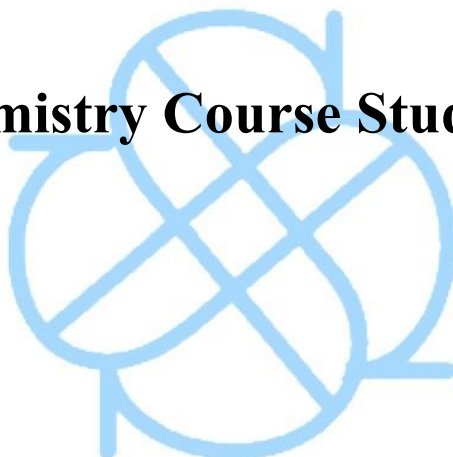


AP Chemistry Course Study Guide



Unit 1: Atomic Structure and Properties - Moles, Mass Spectrometry, Elemental Composition, Mixtures, Atomic

Structure, Electron Configuration, Spectroscopy, and Periodic Trends.

We cannot count particles directly, so we need to find a connection between the mass of a substance and the number of particles. Avogadro's number (6.022×10^{23}) tells us the number of particles in one mole of a pure substance.

The mass of an individual molecule is expressed in atomic mass units (**amu**).

- For example, one oxygen atom has 16 **amu**, but it is also true that one mole of oxygen atoms is equal to 16 grams.

The mass spectrum of a sample containing a single element can be used to determine the identity of the isotopes of that element and the relative abundance of each isotope in nature --AKA how often these isotopes appear in nature.

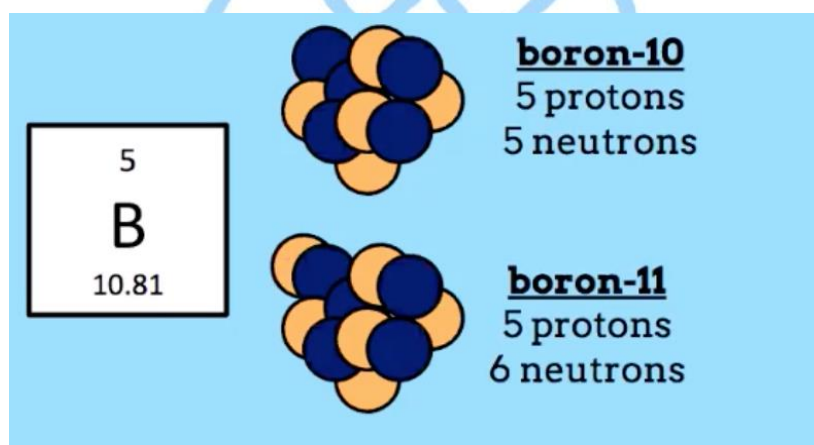


Image from AP Chem YouTube Video from College Board

In order to calculate amu, you take the isotope and multiply it by its percentage.

- Example for boron: $10(0.20) + 11(0.80) = 10.8$ amu

Law of definite proportions: the ratio of mass of constituent elements in any pure sample of that compound is always the same.

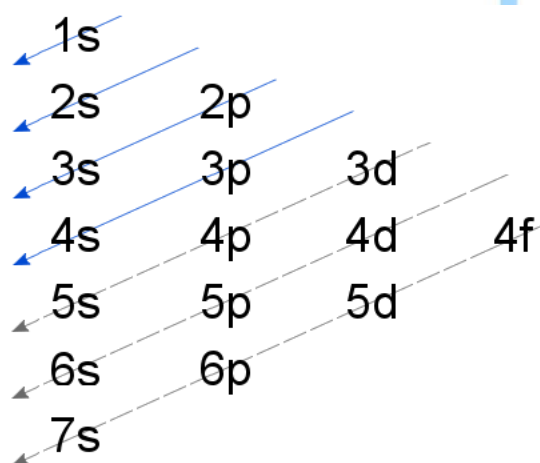
Empirical formula: the lowest and most simplified ratio of elements in a compound. When trying to solve empirical formulas, try to figure out how the moles are related to each other!

A **mixture** is a combination of several molecules all put together. Elemental analysis can be used to determine the relative numbers of atoms in a substance and to determine its purity.

Atoms have negatively charged atoms called **electrons** on the outside and positively charged atoms called **protons** in the nucleus. **Coulomb's law** allows us to calculate the force between charged particles $\rightarrow F = k \frac{Q_1 Q_2}{r^2}$. In atoms and in ions, electrons are organized or thought of as being in **energy levels** (shells) and **sublevels**. This model is described as the electron **configuration model**.

- **Core** electrons are the inner electrons
- **Valence** electrons are outer electrons

Aufbau Principle: electrons fill the lowest energy orbitals (not levels) first.

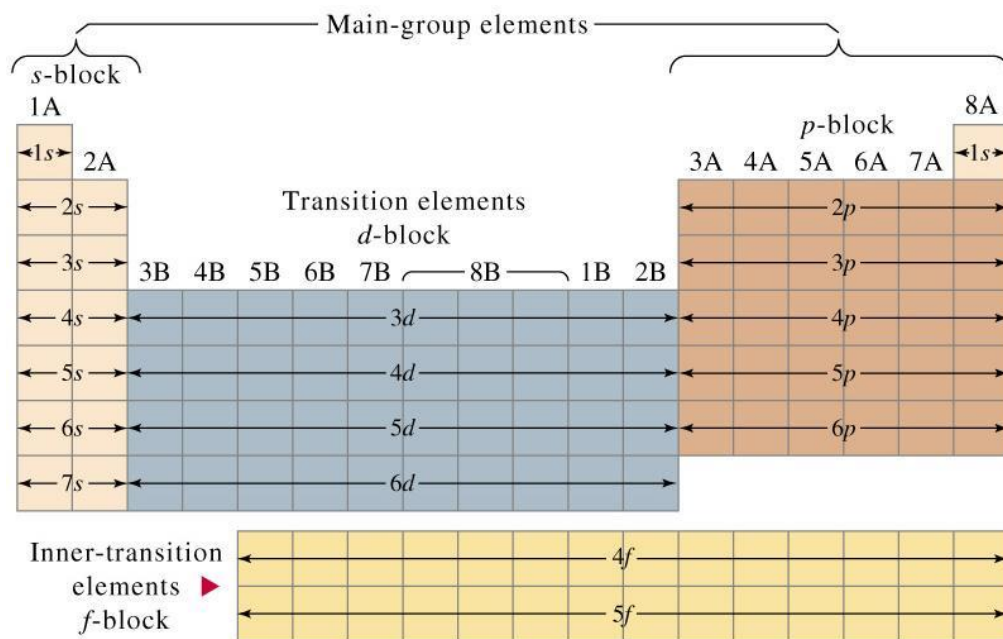


Use this diagram to know which orbitals to fill first. Start from 1s, 2s, then 2p, and so on. (follow the direction of the arrows)

<https://www.chemicool.com/definition/aufbau-principle.html>

With Aufbau principle, it helps us to write out **an electron configuration**:

- For example, Oxygen's configuration would be $1s^2 2s^2 2p^4$



https://www.chem.fsu.edu/chemlab/chm1045/e_config.html

^ This diagram tells us that 2 electrons can be filled in the s orbitals, 6 electrons in the p orbitals, 10 electrons in the d orbitals, and 14 electrons in the f orbitals.

There's a shorthand version for writing electron configuration.

- Ex. Oxygen: $[\text{He}] 2s^2 2p^4$
- Ex. Potassium: $[\text{Ar}] 4s^1$

When you are writing a shorthand version, you can only include a **noble gas** in the bracket.

For ions, you have to remove electrons that are outside leaves first.

- For example, an ion electron configuration for Mg^{2+} would be $[\text{Ne}] 3s^1$
- Ex. Ion electron configuration for Ag: $[\text{Kr}] 4d^{10} 5s^1$.

The 5s orbital lost an electron because it's a more outer sublevel than 4d.

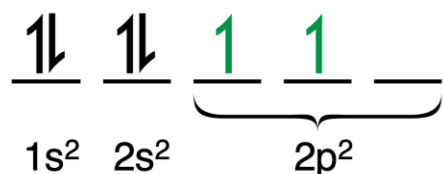
You can check if your electron configuration is right by counting the numbers in the exponent.

- Ex. Remember that the electron configuration for Oxygen is $[\text{He}] 2s^2 2p^4$? In order to check if this configuration is right, you can add 2 and 4, which would give you 6, and this is correct because there are 6 valence electrons for Oxygen.

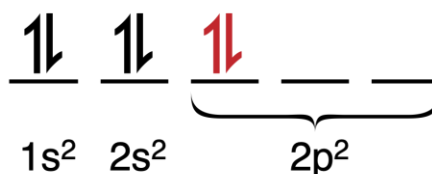
Hund's rule: each sublevel should have one electron before any others are doubled up

- For example, all of the 2p subshells should have at least one arrow in it before moving on to another subshell.

correct

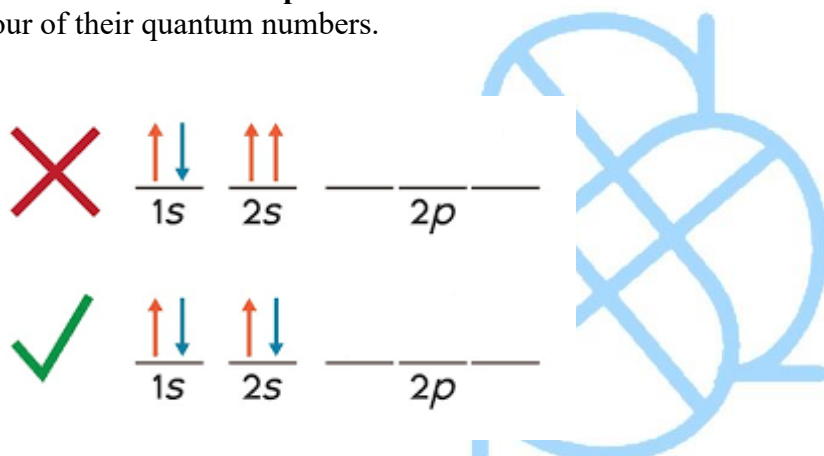


incorrect



<https://ch301.cm.utexas.edu/section2.php?target=atomic/e-config/hunds-rule.html>

Pauli Exclusion Principle: no two electrons in the same atom can have identical values for all four of their quantum numbers.



<https://www.liberaldictionary.com/pauli-exclusion-principle/>

The energies of electrons in a given shell can be measured experimentally with **photoelectron spectroscopy (PES)**. The position of the peak in the spectrum is related to the energy required to remove an electron from the corresponding subshell. The height of the peak is proportional to the number of electrons in that subshell. An electron that is further away from the nucleus will require less energy to break apart.

When reading a PES and trying to determine the electron configuration of an element, start from the left and work towards the right. If you want to double check your answer add up all of the electrons.

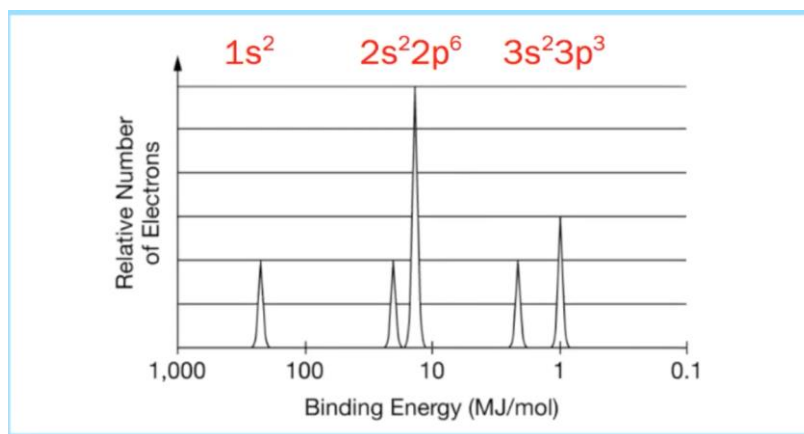


Image from AP Chem YouTube Video from College Board

^ For this graph, assume each horizontal line to be a sublevel. If we look at the first peak, it touches the second horizontal line, then it can be said that that's the $1s^2$. This is similar for the second peak. Now, let's look at the third peak. The third peak touches the 6th horizontal line, then it could be said that it is $2p^6$.

Also, notice how the **binding energy** is the highest for $1s^2$. This means that the electrons have a stronger attraction to the nucleus since it is the closest to the nucleus. In comparison, if we look at the last peak, the binding energy isn't high. This is because the electrons are further away from the nucleus, causing less attraction. Although it was easy to determine the electron configuration from PES, in the actual AP test, it would be more difficult.

The periodic table is organized based on recurring properties. Trends in atomic properties include atomic radius/ionic radius, ionization energy, electron affinity, and electronegativity. Periodicity is used to make predictions in the absence of data.

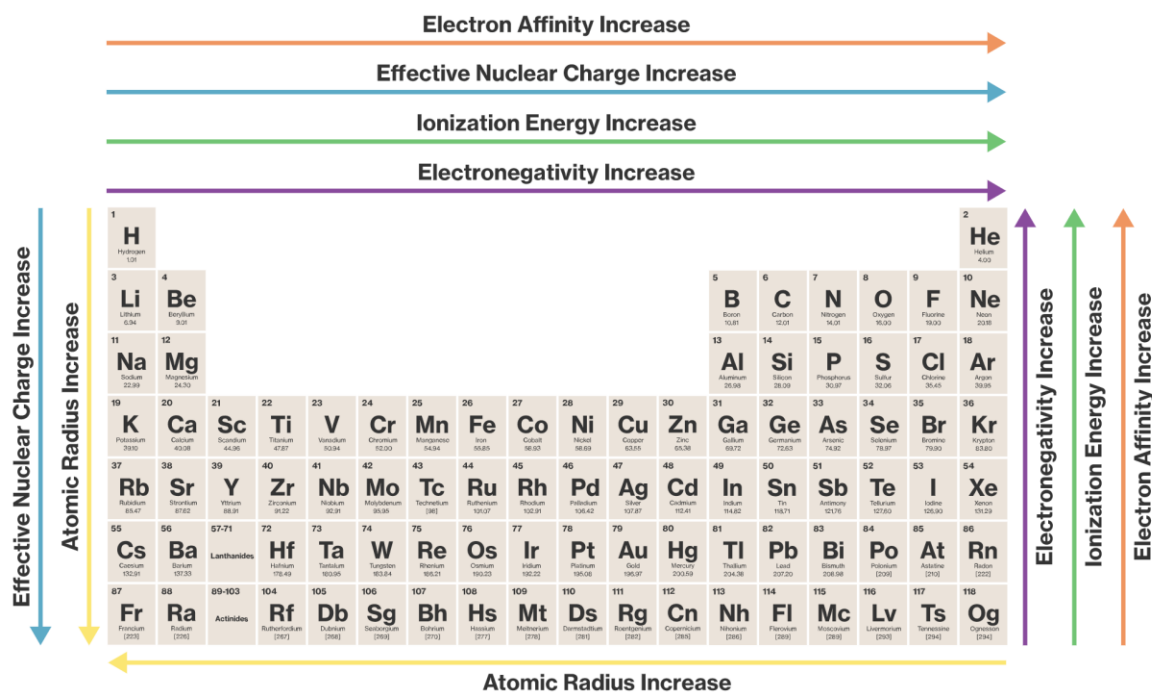
- **Atomic radius:** From left to right on the periodic table, the atomic radius decreases. Although we add protons across the periodic table (protons also considered nuclear charge), the size of the atom will decrease due to **effective nuclear charge**. Effective nuclear charge is a pull toward the charged proton. Since the effective nuclear charge increases, the valence electrons are drawn closer to the nucleus, decreasing the size of an atom. Similarly, atomic radius increases as we go down the table. This is because there are more sublevels added to the atom and due to **shielding effect**. Shielding effect is

when the electron and the nucleus in an atom have a decrease in attraction because the electrons are further away due to increased sublevels.

- For example, Na is bigger than Cl because Na has less effective nuclear charge!
- **Ionic radius:** Cations (positively charged ions) are generally smaller than the atoms that they came from because they have lost valence electrons. The amount of protons always stays the same so this means that since all of those protons are pulling in on the fewer electrons that are there (effective nuclear charge), the atomic radius is significantly smaller.
 - In anions we are adding electrons. There are more electrons on the outside energy levels. Since we have the same number of protons and more electrons, the protons cannot fill as effectively so the radius gets larger.
- **Ionization energy:** Ionization energy is the energy needed to remove an electron. This is strongly associated with atomic radius. Ionization energy increases from left to right because they have a greater effective nuclear charge. Ionization energy decreases as you go down the table because it's easier to remove the electrons due to increased shielding effect.
 - However, there are exceptions: Li has greater Ionization Energy than Be because Li's core electron is being removed.
 - Lowest first ionization energy means the easiest electron to break away from the subshell. This means that you should look for an element with a large atomic radius.
- **Electronegativity:** the ability of an atom to attract electrons toward itself. Electronegativity is a scale. For the AP test, F has the highest electronegativity of 4.0. Electronegativity increases across periods because of increased effective nuclear charge (want more electrons). On the other hand, electronegativity decreases as you go down the table.
- **Electron affinity:** the opposite of electronegativity. It is the energy change when an atom becomes negatively charged. Electron affinity is **always negative**. There isn't a solid trend for electron affinity compared to others. However, this is the general trend: electron affinity becomes **more negative** from left to right. Electron affinity becomes **more positive** when you go down the table.

***On the AP test, remember to talk about the chemistry/science behind it, not the location on the periodic table. This is a very big thing to remember for the test! ***

Here's a chart to memorize these rules easily:



<https://www.coursehero.com/sg/general-chemistry/trends-in-the-periodic-table/>

The likelihood that two elements will form a chemical bond is determined by the valence electrons and nuclei of elements. Elements in the same column/group tend to form similar compounds because of the similar charges.

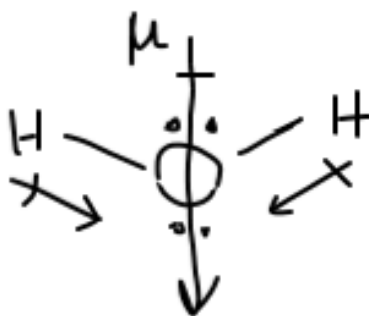
Unit 2: Molecular and Ionic Compound Structures and Properties - Chemical bonds, Intramolecular Force, Structure of Solids, Lewis Structure, Formal Charge, and VSPER.

Let's review basic bonding: Ionic and Covalent

- **Ionic Bonding** is when the electrons are transferred.
- **Covalent Bonding** is when electrons are shared equally.

So... what about intermediate cases?

- **Polar Covalent Bond** is an unequal sharing of electrons between atoms in a molecule and results in a charge separation in the bond. This phenomenon is also called the **dipole moment**. Looking at a water molecule, Oxygen is more negative due to a higher number of electrons. On the other hand, Hydrogen has less electrons. Since Oxygen is more negative, it has higher electronegativity, thus pulling more electrons.



<https://socratic.org/questions/how-do-you-calculate-the-dipole-moment-of-water>

Polarity is dependent on electronegativity.

- Two atoms with the same or similar electronegativity would form a nonpolar bond. The atoms have the same electronegativity so there is no change.
- For example, F-F and Cl-Cl would be non-polar because they have the same electronegativity.
- In a polar bond, one of the atoms will have a larger electronegativity and this causes a pull that makes it polar. The greater the difference in electronegativity the more polar the bond is, like the previous example with water.

When two nonmetals bond together it forms a **covalent** bond. When a nonmetal and a metal bond together it forms an **ionic** bond.

- If a solid is a good conductor of electricity it would have metallic bonding.
- If a solid does not conduct electricity but it becomes a conductor when melted or dissolved in water, then it has ionic bonds.
- If a solid does not conduct electricity in any state (solid liquid, aqueous) and it has a relatively lower melting point, then it has covalent or molecular bonds.
- If a solid does not conduct electricity but has a high melting point, it has a covalent network.

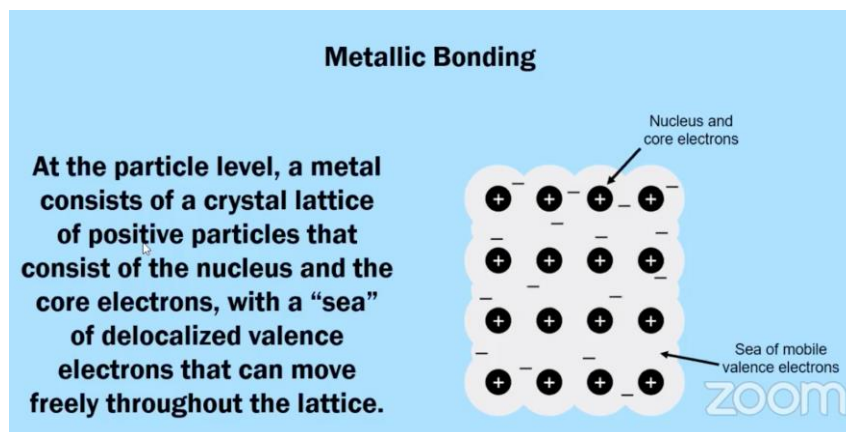
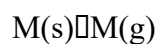


Image from AP Chem YouTube Video from College Board

Lattice energy: the energy required to break apart or separate the ions in a solid crystal lattice (sea of electrons) into individual gaseous ions.

- For the AP exam, it's important to calculate the lattice energy for ionic solid formation.
- Here are the steps for forming an MX solid:

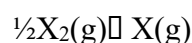
1. Convert to gas for solids:



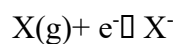
2. Ionization energy of the metal atoms. Depending on the atom, you may need to add 1st Ionization energy 2nd Ionization energy, and so on.



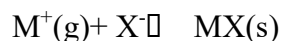
3. Dissociation of the nonmetal if needed



4. Electron Affinity for Formation of X^- ions in the gas phase



5. Lattice Energy for Formation of the Solid MX



Bond Energies: To break bonds, energy must be added to the system and to form bonds, energy is released. The formula for this is

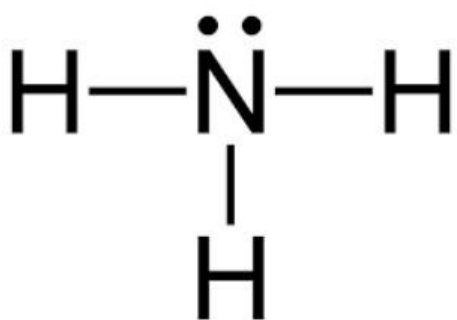
$$\Delta H^\circ_{\text{reaction}} = \sum \Delta H_{\text{bonds broken (reactants)}} - \sum \Delta H_{\text{bonds formed (products)}}$$

<https://study.com/academy/lesson/bond-enthalpy-definition-calculations-values.html>

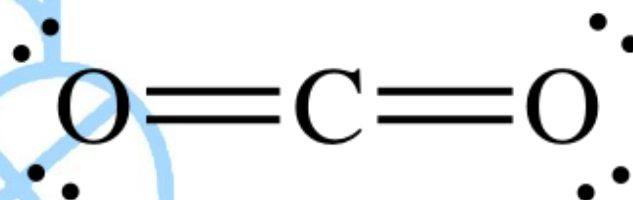
- Bond Energy also depends on the type of bond:
Single Bond < Double Bond < Triple Bond
- The shorter the bond length, the greater the bond energy due to increased attraction between two atoms.

Lewis Structure shows how valence electrons are arranged among atoms in a molecule. There are multiple rules for Lewis Structure:

- **Duet Rule:** This occurs between Hydrogen and Helium and it shares 2 electrons.
- **Octet Rule:** This occurs when each atom has 8 electrons. However, there are exceptions to this rule, which would be discussed later.
- Ex. Lewis Structure for NH_3 :



Ex. Lewis Structure for CO_2 :



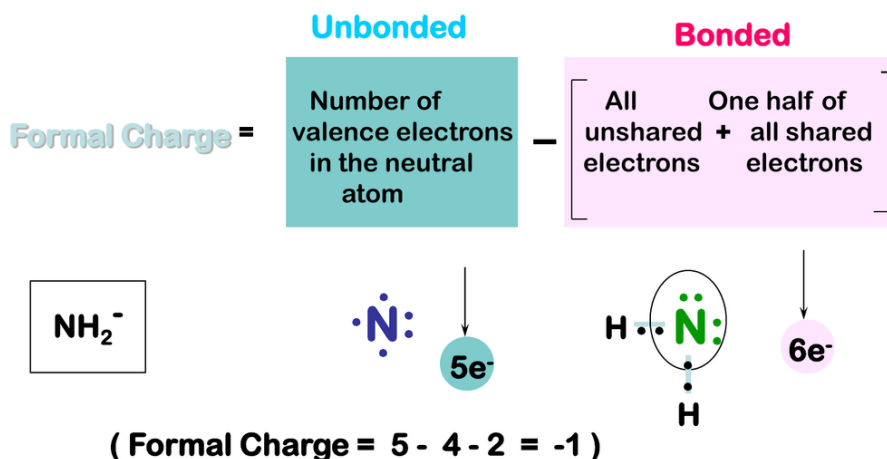
<https://study.com/academy/answer/what-is-the-lewis-structure-of-nh3.html>

<https://socratic.org/questions/how-can-i-draw-a-lewis-dot-diagram-for-carbon-dioxide>

- **Boron** and **Beryllium** form an incomplete octet. Boron can only have 6 electrons, and Beryllium can only have 4 electrons.
- There are also elements that exceed the octet rule. Atoms that exceed row 3 or above, can form an extended octet.
 - For example, Sulfur, Phosphorous, Arsenic can all form an expanded octet.

Formal charge: the charge that you get after subtracting the number of nonbonding electrons and the number of bonds in the molecule from the number of valence electrons in the atom's Lewis structure

FORMAL CHARGE

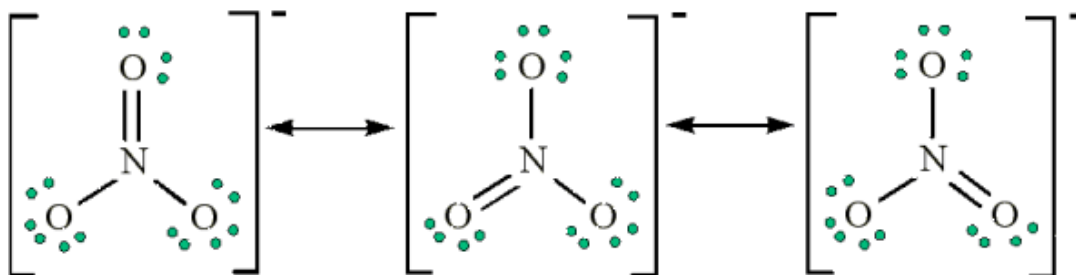


<https://studylib.net/doc/5639925/formal-charge>

- Formal charge is not a real charge; it just helps us to analyze Lewis structures. It's helpful when comparing structures in which each atom obeys the octet rule (like resonance structures). A formal charge of zero is preferred. If it is not possible for each of the atoms to have a formal charge of zero, the more favorable structure is the one where the negative formal charge is on the atom with the highest electronegativity value. Formal charge should also equal to the net charge of a compound.
 - For example, PO_4^{3-} , should have a formal charge of -3 because that's the net charge of a compound.

Resonance structures are when there is more than one way to represent a compound in a Lewis structure. Even though in a resonance structure there is more than one way to represent a compound, technically neither of them is right because in the actual molecule, the bond length is an average of both structures.

- NO_3^- is a good example with 3 resonance structures:



<https://socratic.org/questions/how-would-you-draw-all-the-resonance-structures-for-nitrate-no3>

VSEPR = Valence Shell Electron Pair Repulsion

- When drawing a Lewis structure, we only consider the valence electrons. Electrons in a bonding domain or in a nonbonding domain normally exist in pairs. Electron domains tend to repel each other in a way that is as far apart from each other in 3-D space as possible.
- An electron domain is a bond, but you don't count double or triple bonds as extra domains.
- Lone pairs on the central atom impact the shape.
- Here's an important chart you need to memorize for the AP exam:

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 Pyramidal	 Bent or Angular		
5	 Trigonal Bipyramidal	 Sawhorse or Seesaw	 T-shape	 Linear	
6	 Octahedral	 Square Pyramidal	 Square Planar	 T-shape	 Linear

<https://www.toppr.com/guides/chemistry/chemical-bonding-and-molecular-structure/vsepr-theory/>

Hybridization: Mixing of 2 or more atomic orbitals to form a new set of hybrid orbitals

- sp^3 is a combination of one s and 3 p orbitals, and it is shown on structures with tetrahedral shape. Thus, CCl_4 would have a sp^3 .
- sp^2 is a combination of one s and 2 p orbital, and it is shown on structures with trigonal planar shape. Thus, SO_2 would be an example of this.
- sp is a combination of one s and 1 p orbital, and it is shown on structures with linear shape. Thus, CO_2 would be an example.

Use Geometry of the compound to find hybridization.

- A **sigma** bond is a covalent bond in which the electron density is concentrated in between the two atoms (it is one shared bond).
- A **pi** bond is a covalent bond in which the electron density is concentrated above and below the internuclear axis (double bond).
- A **double bond** represents one sigma bond and one pi bond. A triple bond represents one sigma bond and two pi bonds.

Unit 3: Intermolecular Forces and Properties - IMFs, Solids, Liquids, and Gases, Ideal Gas Law, Kinetic Molecular Theory, Solutions, Mixtures, Solubility, Spectroscopy, Photoelectric Effect, and Beer-Lambert Law.

*Remember that in regards to formal charge, the more favorable structure is the one where the negative formal charge is on the atom with the highest electronegativity!

In a nonpolar molecule: all of the bond dipoles within the molecule cancel each other out. In a polar molecule, all of the bond dipoles within a molecule DO NOT cancel each other out.

A polar molecule tends to be asymmetrical. Nonpolar molecules are symmetrical and that is why the bond dipoles cancel.

Boiling water does not break covalent bonds! When a covalent/molecular substance is heated and undergoes a phase change from a liquid to a gas, that substance has an **intermolecular force**.

****IMPORTANT!!**

<u>Intermolecular (or Interparticle) Forces</u>	
The attractive forces that exist IN BETWEEN MOLECULES (or particles)	WEAK particles are relatively far apart
<u>Intramolecular Forces</u>	
The attractive forces that exist WITHIN A MOLECULE	STRONG particles are relatively close together

Image from AP Chem YouTube Video from College Board

London dispersion forces (or Van der Waals forces) occur when the instantaneous dipole (caused by moving electrons in an individual atom) in one atom can cause another atom to have an uneven distribution in its electron cloud.

- LDFs are all about the electron cloud!
- **A larger electron cloud has greater polarizability which means that it can be easily distorted (more room for electrons to be spontaneous= larger instantaneous dipole). The larger the electron cloud, the stronger the LDF. All molecules experience these.**
 - On the exam never say that an atom has stronger LDFs because it is heavier!!
Talk about the larger electron cloud.

Polar molecules have stronger IMFs.

The attractions between POLAR molecules are called **dipole-dipole** interactions. This is when the partial positive end is attracted to the partial negative end of another molecule.

The polar molecule will NOT ALWAYS have the stronger IMF. Always remember to consider the electron cloud (for example: CCL_2 vs. HCl - CCl_2 . HCl - CCl_2 is more polar because of its larger electron cloud).

Hydrogen Bonding is a strong dipole-dipole interaction that occurs between a Hydrogen atom and either F, O, or N (H-F, H-O, and H-N)

Ion-dipole forces are INTERPARTICLE. This is the attractive force between an ion and a polar molecule.

Stronger intermolecular forces have **higher** melting point and boiling point.

For the AP exam, it is important to know the characteristics of Solids, Liquids, and Gases.

First, liquid has a low compressibility, lack of rigidity, and high density compared to gas.

- Liquids with large intermolecular forces have **high surface tensions**
 - **Surface tension:** resistance of a liquid to an increase in its surface area
- Liquids with large intermolecular forces or molecular complexity tend to be viscous.

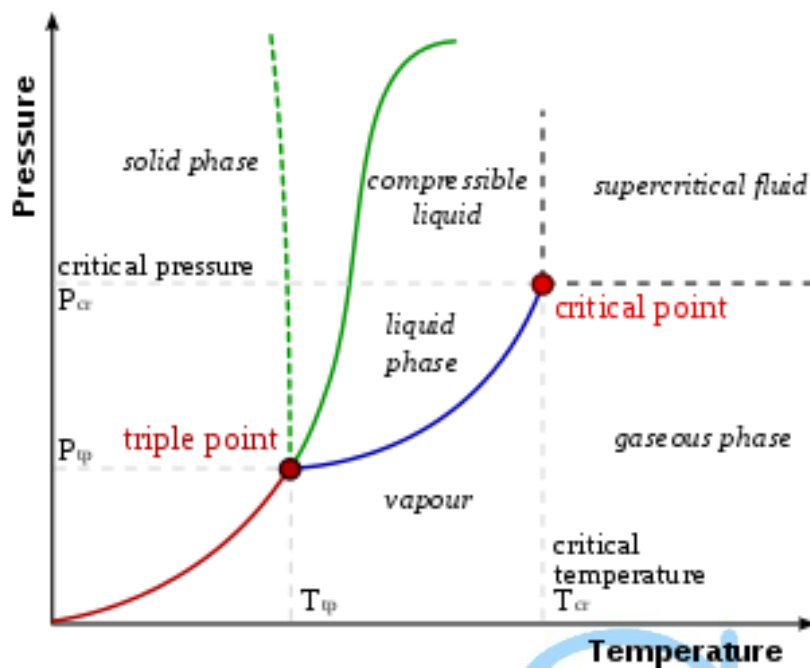
Solids:

- **Crystalline Solid:** Rigid, and long-range order. Predictable positions.
- **Amorphous Solid:** Does not possess a well-defined arrangement or long-range molecular order (usually like a powder)

Vapor pressure is the pressure exerted on a gas at equilibrium with its liquid phase at a given temperature. (l) \rightarrow (g). **Stronger** intermolecular forces have **less vapor pressure** because it sticks to other molecules.

- Warming a liquid raises the vapor pressure because a greater fraction of the molecules have the **Kinetic Energy** needed to escape the liquid; this increases the vapor pressure because it will cause more gas particles in the container.
- **Lower** boiling point means **higher** vapor pressure because that means it will take less heat energy to turn a liquid into a gas. Since it will take less energy, this increases the vapor pressure because the molecules will have more KE to turn Liquid \rightarrow Solid

Phase Diagrams are a convenient way of representing phases of substance as a function of temperature and pressure.



Triple Point is when all the lines meet up.

Critical Point is when a gas can't be made to liquefy no matter how great the pressure and temperature is.

https://en.wikipedia.org/wiki/Phase_diagram

Gas:

- Uniformly fills any container.
- Mixes completely with any other gas
- Exerts pressure on its surroundings.

For the AP exam, there are many important laws to memorize.

Boyle's Law: Pressure and volume are inversely related under the condition that temperature is constant.

$$P_1V_1 = P_2V_2$$

P_1, P_2 are pressure of gas

V_1, V_2 are volume of gas

<http://matkarimbo.blogspot.com/2014/07/gas-laws.html>

Charles's Law: Volume and temperature in Kelvin are directly related under the condition that pressure is constant.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

T = Temperature of the gas

V = Volume of the gas

Pressure must be constant

<https://sites.google.com/site/tiyahgaslaws/charles-s-law>

Gay-Lussac's Law: Pressure and Temperature are directly proportional if volume is constant.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

<https://introductiontogaslaws101.weebly.com/gay-lussacs-law.html>

Combined Gas Law:

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

https://www.chem.fsu.edu/chemlab/chm1045/gas_laws.html

Avogadro's Law: Volume is directly related to the number of moles (represented as n) under the condition that temperature and pressure is constant.

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

https://www.chem.fsu.edu/chemlab/chm1045/gas_laws.html

If we combine all these laws, we get the **Ideal Gas Law:**

PV=nRT

**R = gas constant = 0.08206 L*atm mol⁻¹ K⁻¹

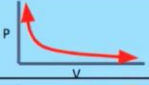
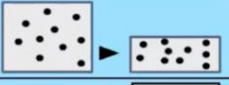
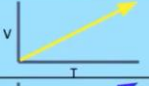
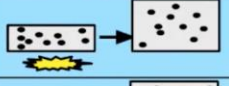
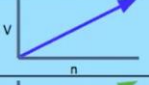
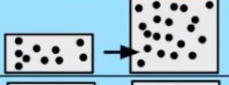
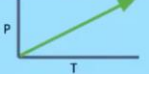

3.4: Ideal Gas Law			
Variables Gas Law	Graphical Representation	Picture	Example
Pressure vs volume (Boyle's Law)			If I push on a balloon (decrease the volume) the pressure inside the balloon increases.
Volume vs Temperature (Charles' Law)			A balloon decreases in volume if you put it outside in the cold .
Volume vs Moles (Avogadro's Law)			A balloon increases in volume as you add more (moles of) air to it.
Pressure vs Temperature (Gay-Lussac's Law)			When the temperature of a container increases, the pressure of the gas inside increases also.

Image from AP Chem YouTube Video from College Board

STP is standard pressure at 1 atm and standard temperature of 273 K. Also, under STP, 1 mol equals to 22.4 L.

Dalton's Law of Partial Pressures:

Dalton's Law of Partial Pressures

For a mixture of gases in a container,

$$P_{\text{Total}} = P_1 + P_2 + P_3 + \dots$$

<https://www.thinglink.com/scene/508702545376444417>

Kinetic Molecular theory:

3.5: Kinetic Molecular Theory

The kinetic molecular theory (KMT) tells us WHY ideal gases behave as they do. There are 5 "postulates":

- 1) Gases are composed of molecules that are in continuous motion, travelling in straight lines and changing direction only when they **collide** with other molecules or with the walls of a container.
- 2) The pressure exerted by a gas in a container results from **collisions** between the gas molecules and the container walls.
- 3) The **average kinetic energy** of the gas molecules is proportional to the kelvin temperature of the gas.
- 4) The molecules composing the gas are **negligibly small** compared to the distances between them.
- 5) Gas molecules exert **no attractive or repulsive forces** on each other or the container walls; therefore, their collisions are elastic (do not involve a loss of energy).

Image from AP Chem YouTube Video from College Board

The Ideal gases law does not represent real gases. They are more examples of how gases would act in an ideal/perfect space. Deviations from the ideal gas law may result from:

- Interparticle attractions amongst gas molecules, particularly those close to condensation
- Extremely high volumes
- At high pressures (low temp and volume) the volumes of the particles are no longer negligible compared to the volume of the container.
- Some gas particles DO attract or repel one another. These attractive or repulsive forces can impact the strengths of these collisions.

In a homogeneous mixture the macroscopic properties do not vary throughout the sample.

In a heterogeneous mixture, the macroscopic properties really depend on that specific location in the mixture.

Substances with similar intermolecular interactions tend to be soluble with one another. Same with polar and nonpolar substances.

“Like dissolve like”

Root Mean Square Velocity

$$u = \sqrt{\frac{3RT}{M}}$$

M = molar mass of gas in kg/mol.

$R = 8.3145 \text{ J/K} \cdot \text{mol}$.

Final Unit is in m/s

<https://dvhsgaslaws.weebly.com/root-mean-square-velocity.html>

Diffusion: the mixing of gas

Effusion: The passage of a gas through a tiny orifice into an evacuated chamber.

Graham's Law of Effusion:

$$\frac{R_1}{R_2} = \sqrt{\frac{M_2}{M_1}}$$

R represents the rate of effusion for each gasses.

<https://calistry.org/calculate/grahamLaw>

Unit 4: Chemical Reactions - Reactions, Net Ionic Equations, Chemical Changes, Stoichiometry, Titrations, Acid-Base Reactions, and Redox Reactions.

A **physical change** occurs when a substance undergoes a change in properties, but not a change in its chemical composition.

- For example, phase changes (solid to liquid, liquid to gas, etc.).

A **chemical change** occurs when substances are transformed into new substances, with different compositions.

- For example, the combination of CO_2 and H_2 to form ethanol CH_3OH .

Make sure to balance all reactions, especially net ionic ones!

When writing net ionic reactions, know that water is not a reactant. It is a medium used to dissolve the solid so it does not need to be written in the equation.

A (covalent) molecular substance that is soluble in water, like sugar, but does not react with water, will exist in the solution at the particle level of molecules. It does not dissociate into ions.

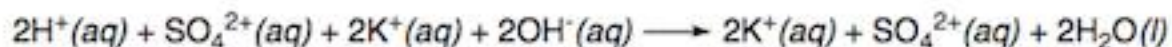
How to write Net ionic equations:

- Start with molecular equation
- Make the transition to the complete ionic equation
- Cross off spectator ions
 - Spectator ions are ions that appear on both sides of the equation and they can be canceled out.
- Here's an example:

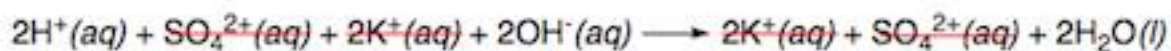
Balanced reaction equation.



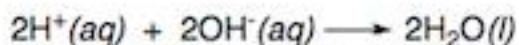
Ionic equation.



Cross out the spectator ions.



Net ionic equation.

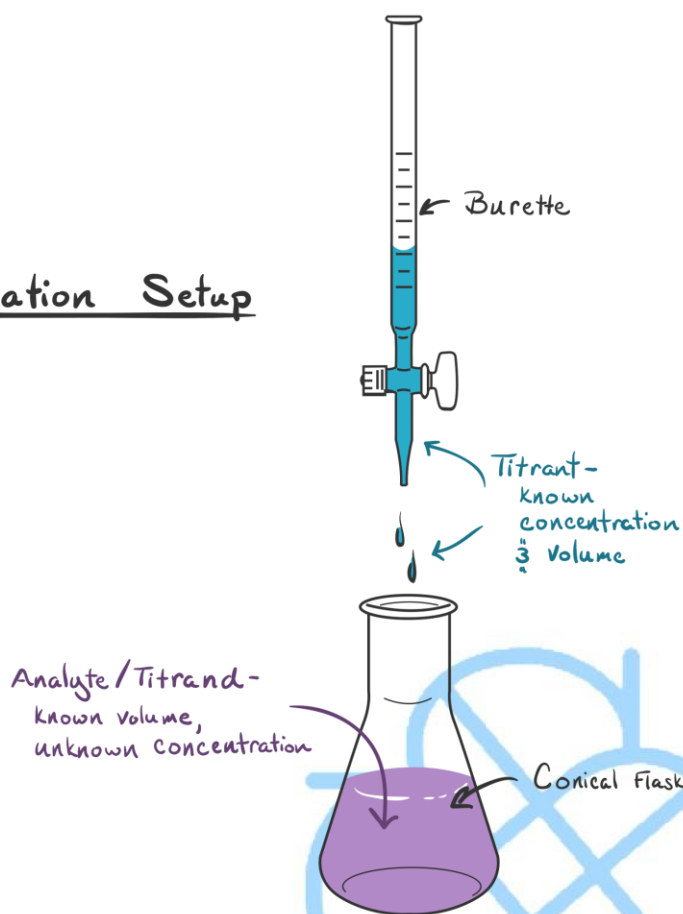


<https://study.com/academy/answer/write-the-net-ionic-equation-for-the-reaction-between-sulfuric-acid-and-potassium-hydroxide.html>

A **chemical formula** should be written as is if it is a strong acid, a strong base, or a soluble ionic compound. When this happens, split it up into its ions. It should be split up into separate ions if it is a covalent molecular substance, a weak acid or a weak base, a gas, a pure liquid, a pure solid, or an insoluble ionic solid.

A **titration** is an experiment that can be used to calculate the amount of moles or the concentration of one chemical by reacting it with a known quantity of another chemical. The standard solution of the known concentration is known as the titrant. The solution of unknown concentration is known as the **analyte**.

Titration Setup



<https://www.khanacademy.org/test-prep/mcat/chemical-processes/titrations-and-solubility-equilibria/a/acid-base-titration-curves>

A bronsted lowry acid is a proton donor (H^+) and a bronsted lowry base is a proton (H^+) acceptor.

Oxidation redox reactions involve the transfer of electrons from one substance to another. The species that LOSES electrons is oxidizing. The species that GAINS electrons is reducing.

The oxidation number will go up in number when it oxidizes and down in number when it is reducing.

There is mnemonic device to memorize this: **LEO the Lion says GER**



<https://me.me/i/leo-the-lion-says-ge-r-lose-electrons-oxidation-gain-48ca5351a133405bae12633078f25e18>

These equations usually have the pattern of having an element in the reactant that turns into a product on the products side, and vice versa.

Oxidation number rules:

- An atom in its elemental form is always zero.
- For a monatomic ion, the oxidation number is equal to the charge on the ion.
- The oxidation number of fluorine is always -1
- The oxidation number of other halogens is almost always -1 (Cl, Br, I). When those atoms are bonded to fluorine or oxygen, it will have a positive oxidation number.
- Hydrogen is normally +1 when bonded to other nonmetals. When bonded to a metal it is -1
- Oxygen is usually -2. However, in peroxides the oxidation number for oxygen is -1.

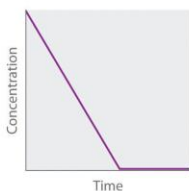
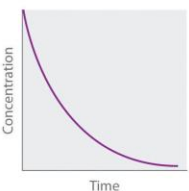
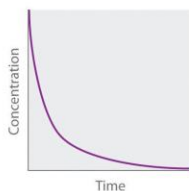
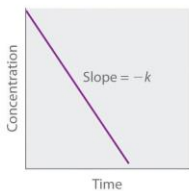
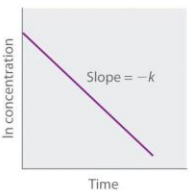
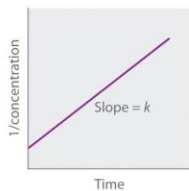
In order to prove if a reaction is oxidizing or reducing you NEED to break the equation down and assign oxidation numbers so that you can clearly see the change.

The oxidation half reaction shows the loss of electrons. The reduction half reaction shows the gain of electrons.

Unit 5: Kinetics - Reaction rates, Rate law, Concentration changes, Reaction mechanisms, Steady State approximation, Collision Models, Reaction energy profiles, and Catalysis.

Chemical Kinetics refers to the rates of chemical reactions and the mechanisms at which they occur. The rate of a reaction is the change in concentration of a product over a unit of time. The rate and concentration of a substance are directly related. The unit of K will not be the same for each reaction.

Here's a chart that shows the graphs, differential rate law, integrated rate law, half-life, and units for each order: Zero, First, and Second. It is important that you memorize these concepts for the AP exam.

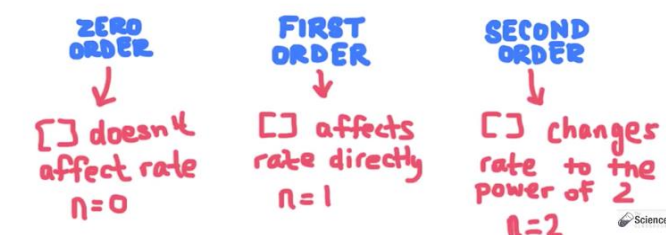
	Zeroth Order	First Order	Second Order																								
Differential rate law	Rate = $-\frac{\Delta[A]}{\Delta t} = k$	Rate = $-\frac{\Delta[A]}{\Delta t} = k[A]$	Rate = $-\frac{\Delta[A]}{\Delta t} = k[A]^2$																								
Concentration vs. time																											
Integrated rate law	$[A] = [A]_0 - kt$	$[A] = [A]_0 e^{-kt}$ or $\ln[A] = \ln[A]_0 - kt$	$\frac{1}{[A]} = \frac{1}{[A]_0} + kt$																								
Straight-line plot to determine rate constant																											
Relative rate vs. concentration	<table><tr><th>[A], M</th><th>Rate, M/s</th></tr><tr><td>1</td><td>1</td></tr><tr><td>2</td><td>1</td></tr><tr><td>3</td><td>1</td></tr></table>	[A], M	Rate, M/s	1	1	2	1	3	1	<table><tr><th>[A], M</th><th>Rate, M/s</th></tr><tr><td>1</td><td>1</td></tr><tr><td>2</td><td>2</td></tr><tr><td>3</td><td>3</td></tr></table>	[A], M	Rate, M/s	1	1	2	2	3	3	<table><tr><th>[A], M</th><th>Rate, M/s</th></tr><tr><td>1</td><td>1</td></tr><tr><td>2</td><td>4</td></tr><tr><td>3</td><td>9</td></tr></table>	[A], M	Rate, M/s	1	1	2	4	3	9
[A], M	Rate, M/s																										
1	1																										
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Half-life	$t_{1/2} = \frac{[A]_0}{2k}$	$t_{1/2} = \frac{0.693}{k}$	$t_{1/2} = \frac{1}{k[A]_0}$																								
Units of k, rate constant	M/s	1/s	M ⁻¹ ·s ⁻¹																								

<https://2012books.lardbucket.org/books/principles-of-general-chemistry-v1.0/s18-04-using-graphs-to-determine-rate.html>

Rate law: mathematical expression that shows how the rate of a reaction depends on the concentration of the reactants.

REACTION ORDER

DETERMINES HOW [REACTANTS]
AFFECTS REACTION RATE



- The rate law includes the reactants ONLY. The exponents above the reactants when combined will reveal the order of the reaction.

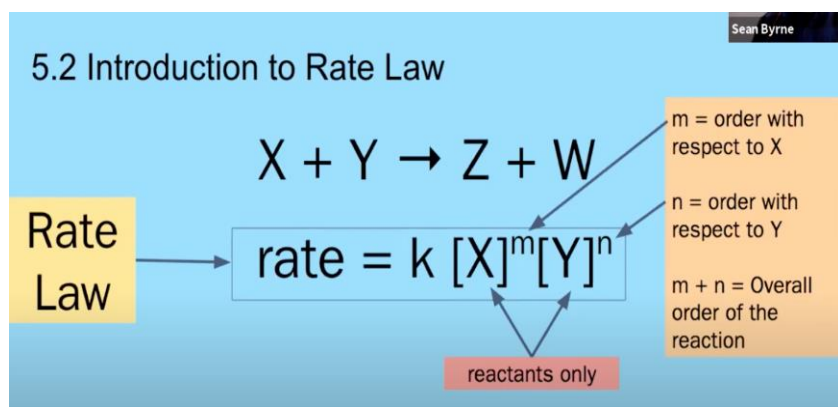


Image from AP Chem YouTube Video from College Board

If the change is the same for both of the concentrations and the rate, it is considered a first order reaction because there is a direct relationship. In the chart if you notice that both the rate and the concentration doubles it is likely a first order reaction.

Integrated rate laws relate the concentration of a reaction to the time

- *This differs from a rate law because it relates rate laws of a reaction to the concentration of the reaction.

A rate is known as the speed at which a reaction proceeds in a certain direction. (Rate=m/s) **Rate** = $K [A][B]$. The rate constant (K) defines how fast the rate goes. It is at a specific temperature. The unit of K reflects the overall order and has variable units.

The **reaction rate** is the measure of how quickly the reactants in a reaction change into the products of the reaction. It can be measured by how quickly the reactants decrease and the second way is to measure how quickly the products increase.

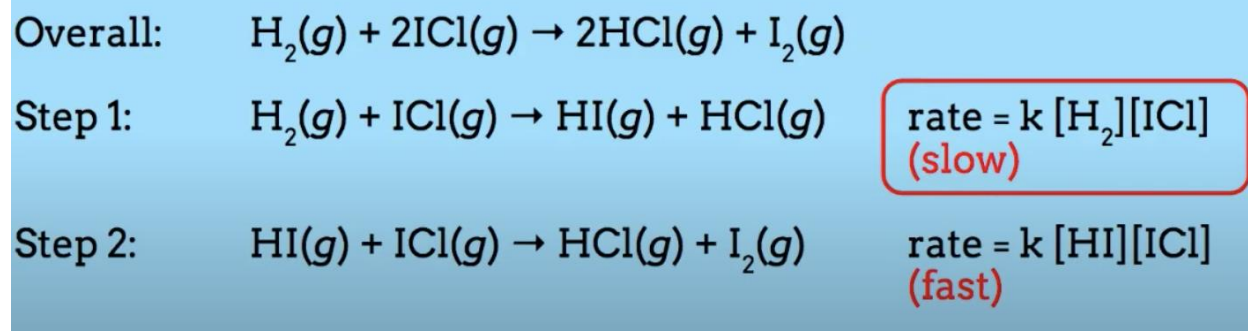
Reaction mechanisms - An elementary reaction is a normal balanced equation. When these equations are broken down, it can help us to find the reaction rate during different parts of the reaction. This can help when determining the slow and fast states of a reaction. The slower reaction is the rate determining reaction.

In a reaction, there are:

- **reactants** - substances that only appear on the left hand side of a step
- **Products** - substances that only appear on the right hand side of a step
- **Intermediates** - appears as a product in an earlier step and then as a reactant in a later step
- **Catalysts** - these interact with the reactant in the first step and are produced again as a product so it doesn't change the reaction.

*Intermediates are not included in the rate law!!!

Not all chemical reactions can occur in one step, so we must look at the series of steps. Each step in a mechanism is called an **elementary reaction**. Species that are in elementary reactions but are not in the final overall equation are called intermediates (they are in the products in one equation and then appear as reactants in another- this causes them to cancel). Whenever dealing with mechanisms, the **slowest** step is **the rate determining step** - the rate of that particular step determines the rate of the entire reaction. Remember, when there is more of a substance in a reaction, that will increase the rate of the reaction because it will take longer for the reaction to be complete.



<https://2012books.lardbucket.org/books/principles-of-general-chemistry-v1.0/s18-04-using-graphs-to-determine-rate.html>

In the image above, the correct kinetics equation would be $\text{rate} = k[\text{H}_2][\text{ICl}]$ because step 1 was the rate-determining step.

Collision Model: In a reaction, the particles in the reactants must successfully collide in order to produce products. The bonds need to be broken.

- Successful collisions have the right amount of energy and the correct orientation that allows the bonds to rearrange in the required manner.
- The **Collision theory** states that, for a reaction to occur, particles must collide with the correct orientation and with sufficient energy for a reaction to occur. Different factors affect the rate of the reaction by affecting the frequency of particle collisions, and/or the proportion of collisions that have enough energy to react.
- More particles (higher concentration) = more opportunities for collisions to happen.
- Increasing temperature increases the frequency of a reaction and the % of a successful collision.
- Increasing the surface area creates more opportunities for collisions. This however does not mean that the energy or the strength has increased, it just increases the likelihood.

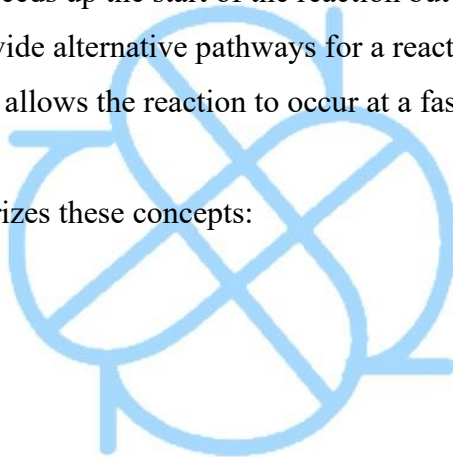
The main factors that affect reaction rate include:

- **Reactant concentrations:** The concentration can be increased by increasing the number of particles or by decreasing the volume and keeping the number of particles the same. In a diluted solution, there are not many particles to react in a given time, causing a slow reaction. In a more concentrated solution, there are more particles in a given volume that

causes a faster reaction because it requires less time to react. This also causes a higher reaction rate.

- **Pressure:** If pressure increases, volume will decrease. This decrease in volume causes an increase in concentration. This causes more collisions and leads to an increase in the rate of the reaction.
- **Temperature:** Particles need a certain amount of energy (activation energy) for collisions to be successful. Increasing the temp increases the average kinetic energy of the particles. More of them will have the energy to react and this will lead to an increase in the rate of the reaction.
- **Surface Area:** When surface area is increased, the reaction rate increases because the particles in a solution can react faster.
- **Catalyst:** A catalyst speeds up the start of the reaction but is not actually used up in the reaction. Catalysts provide alternative pathways for a reaction to occur that takes up less activation energy. This allows the reaction to occur at a faster rate.

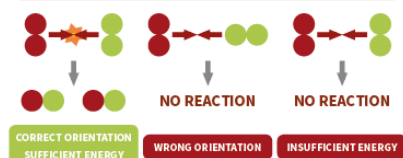
Here's a chart that summarizes these concepts:



MAKING CHEMICAL REACTIONS HAPPEN FASTER

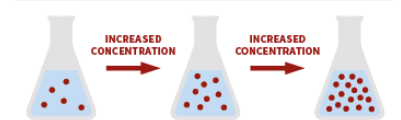
There are a number of different things that we can change to make a chemical reaction faster. Here, we explain the concept of collision theory, and how it can be used to explain the effects of five different factors on the rate of a chemical reaction.

COLLISION THEORY



Collision theory states that, for a reaction to occur, particles must collide with the correct orientation and with sufficient energy for a reaction to occur. Different factors affect the rate of the reaction by affecting the frequency of particle collisions, and/or the proportion of collisions that have enough energy to react.

INCREASE CONCENTRATION OF REACTANTS

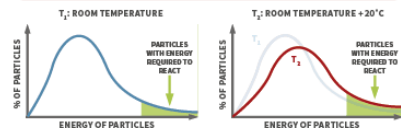


↑ FREQUENCY OF COLLISIONS

— % SUCCESSFUL COLLISIONS

Increasing the concentration of reactants in solution increases the rate of reaction as there are a greater number of particles available to react. This increases the frequency of collisions between particles.

INCREASE TEMPERATURE OF REACTION



Increasing the temperature increases the kinetic energy of particles. This increases the frequency of particle collisions, and a greater proportion of collisions will have the energy required to react.

INCREASE SURFACE AREA OF REACTANTS

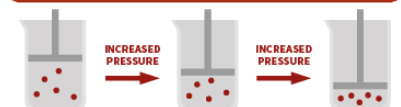


↑ FREQUENCY OF COLLISIONS

— % SUCCESSFUL COLLISIONS

Increasing the surface area of solid reactants increases the number of particles that are exposed and available to react, and as a consequence this increases the frequency of particle collisions, increasing rate.

INCREASE PRESSURE OF REACTION

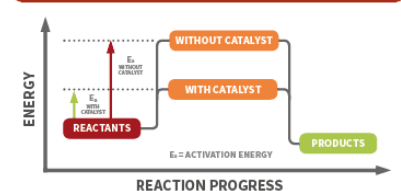


↑ FREQUENCY OF COLLISIONS

— % SUCCESSFUL COLLISIONS

Increasing the pressure of a reaction involving gases forces the gas particles closer together. This will increase the frequency of particle collisions, and therefore increase the rate of reaction.

USE A CATALYST IN THE REACTION



A catalyst provides an alternative route for the reaction, with a lower activation energy. This means that particle collisions need less energy in order for a reaction to occur, increasing the rate of the reaction.



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<https://i1.wp.com/www.compoundchem.com/wp-content/uploads/2016/02/Factors-Affecting-Rate-of-Reaction.png?ssl=1>

Unit 6: Thermodynamics - Energy Diagrams, Thermal Equilibrium, Heat Capacity, Enthalpy of Reaction/Formation, and Hess's Law.

In an endothermic process, energy (in the form of heat) is transferred from the surroundings to the system, taking in energy.

- The system is the reaction - the thing that is being dissolved or reacting. The surroundings is everything outside of the surroundings like the air or container.

In an exothermic process, energy is transferred from the system to the surroundings, pushing energy out.

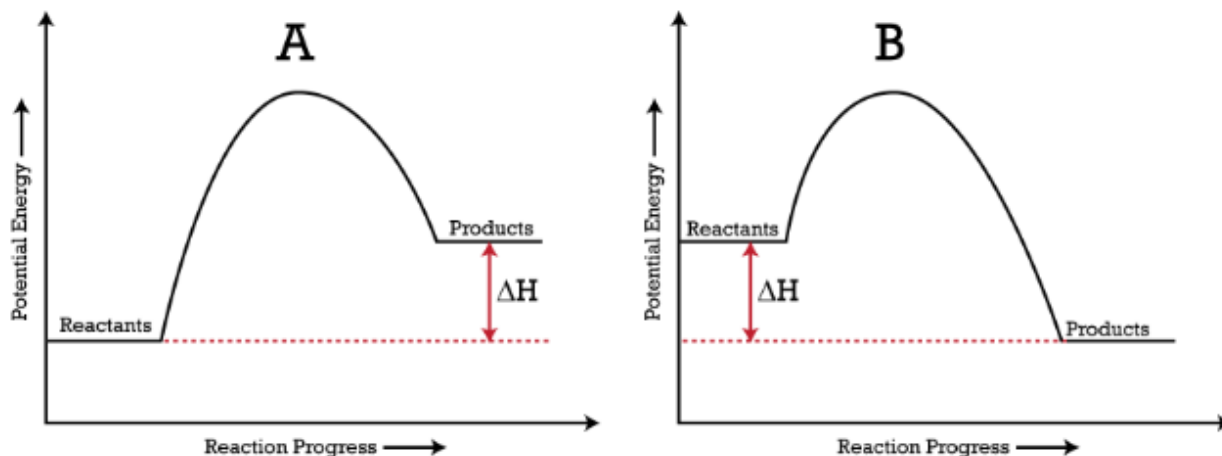
The system is the portion of the experiment that we choose to focus on. It normally consists of the reactants and products. The surroundings are everything else such as the container or everything outside.

The **law of thermodynamics** states that heat cannot be created or destroyed but instead can be transferred from one form to another.

- In an endothermic reaction, heat is **ABSORBED** by the system. This means that work is done on the **SYSTEM**.
- In an exothermic reaction, heat is **RELEASED** into the surroundings. This means that work is done on the **SURROUNDINGS**.

In an energy diagram you can tell if the reaction is endo or exo based off of the placement of the reactants and products.

- In endothermic reactions, the reactants start low and products end high (see Graph A)
- In exothermic reactions, the reactants start high and products end low because energy is being released (see Graph B)



[https://chem.libretexts.org/Bookshelves/Introductory_Chemistry/Book%3A_Introductory_Chemistry_\(CK-12\)/18%3A_Kinetics/18.4%3A_Potential_Energy_Diagrams](https://chem.libretexts.org/Bookshelves/Introductory_Chemistry/Book%3A_Introductory_Chemistry_(CK-12)/18%3A_Kinetics/18.4%3A_Potential_Energy_Diagrams)

Energy is the capacity to do work or transfer heat.

Heat is the transfer of energy that takes place because of a difference in temperature. Heat spontaneously transfers from warm to cool objects.

Work is the exertion of a force through a distance.

Enthalpy is used to measure changes in a constant-pressure system. When we are talking about enthalpy (ΔH), we are basically talking about heat.

- If enthalpy is **positive** it's an **endothermic** process
- If enthalpy is **negative** it's an **exothermic** process.

When it comes to Limiting reactant, the excess is always the element with the higher amount of moles!! The limiting reactant will generally have the lower number of moles by comparison.

Bond enthalpy is the energy required to break a bond. It can also represent the amount of energy released when a bond is formed. ($\sum \text{Bonds broken} - \sum \text{Bonds Formed}$).

Enthalpy of formation is **products-reactants**.

Hess's Law is an enthalpy change for the overall process and is a sum of enthalpy changes of each step in the process.

- Equation: $\Delta H_{\text{rxn}} = \Delta H_1 + \Delta H_2 + \Delta H_3 + \dots$

Unit 7: Equilibrium - Equilibrium, Reversible Reactions and the Equilibrium Constant.

Many think that when equilibrium occurs, there are equal amounts of reactants to products. That is a BIG misconception!

The time at which equilibrium is met means that the concentrations of the forward and reverse reactions remain **constant**/do not change. When equilibrium is met there are no observable changes occurring in the system.

When writing equilibrium equations, the general rule is products over reactants!

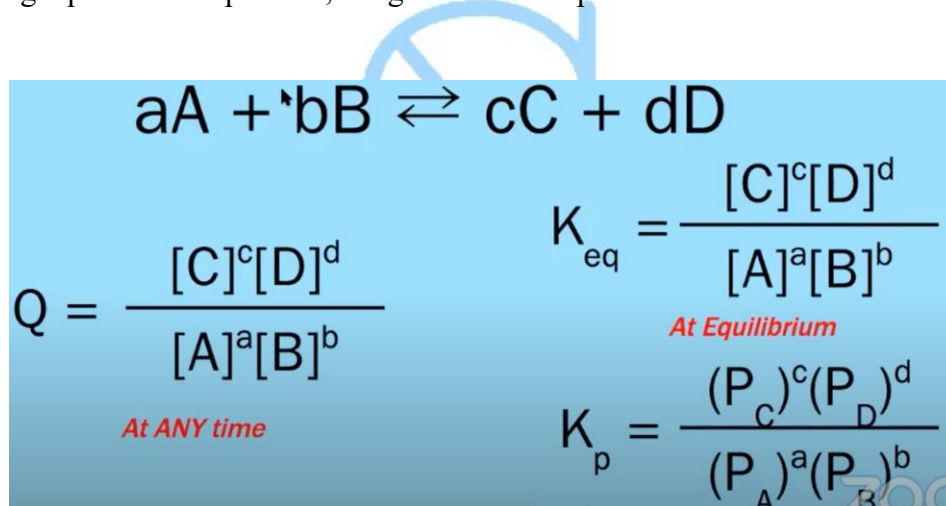


Image from AP Chem YouTube Video from College Board

The **reaction quotient** (Q) measures the relative amounts of products and reactants present during a **reaction** at a particular point in time. The **reaction quotient** aids in figuring out which direction a **reaction** is likely to proceed, given either the pressures or the concentrations of the reactants and the products. **K** is the constant of a certain reaction when it is in equilibrium, while **Q** is the quotient of activities of products and reactants at any stage of a reaction.

- If $Q > K$, the system will proceed in the **reverse** direction. If $Q < K$, the system will proceed in the direction of the **forward** reaction.
- The reaction quotient normally does NOT include solids or pure liquids.

- The equilibrium constant for a gas should always be in parenthesis, but the equilibrium constant for anything else should be in brackets to indicate concentration.

In an **ICE** chart, the -x is always the known value of moles. If you don't have a concentration value or that value starts at zero, it will be +x

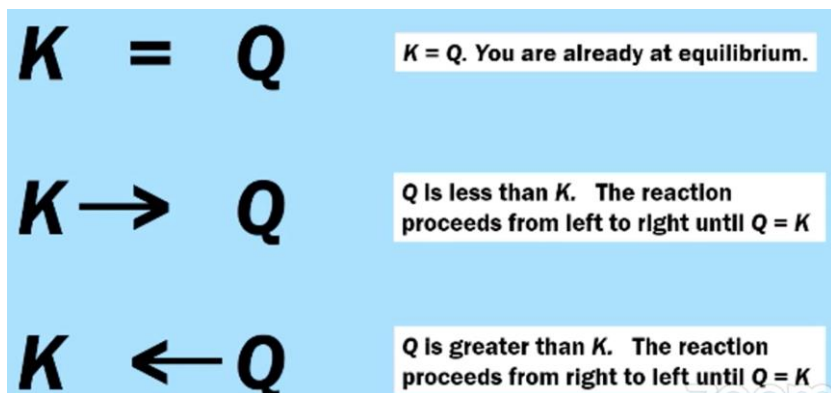


Image from AP Chem YouTube Video from College Board

If a system at equilibrium experiences some sort of stress, the equilibrium shifts in the direction that relieves that stress.

If K is a large number, it means that the equilibrium concentration of the products is large. In this case, the reaction as written will proceed to the right (resulting in an increase in the concentration of products)

- Reactions with a large K value usually go to completion

The larger the **K_{sp}** value, the larger the solubility would be (for most 1 to 1 ratios).

The **common ion effect** states that the solubility of an ionic solid is going to be less/reduced when it is dissolved into a solution that already contains one of the ions present in the ionic solid.

An ionic solution is more soluble in an acidic solution.

When predicting the production of a precipitate, remember that when you add mL or more water to the solution, the concentration of ions goes down/decreases. In order to determine if a precipitate will form, do a Q vs. K comparison.

We calculate Q based on the concentration of each ion in the final solution.

- If $Q > K_{sp}$, a precipitate forms.
- If $Q < K_{sp}$, no precipitate forms.

SHORT REVIEW OF EQUILIBRIUM

- **Le Chateliers Principle:** if you induce stress on a system, the equilibrium will shift to reduce that stress.
- Changing the concentration, changing the temperature (higher temperature means more heat, so heat will shift to the other side to relieve the stress on the other side), and changing the volume and pressure will cause a shift in the equilibrium.
- Here's a chart for easier understanding:

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.

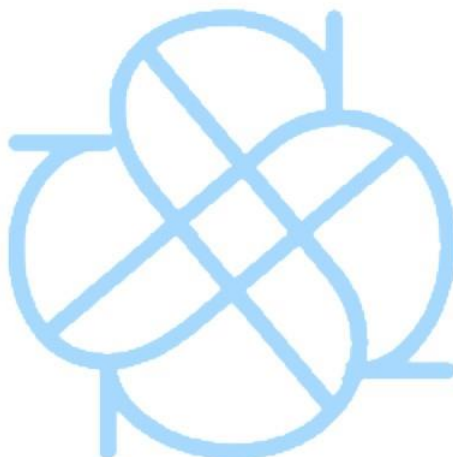
Important information:

when there's a change in temperature, **K Changes**. However, for other changes, K remains the same.

<https://www.pinterest.com/pin/19773685846742017/>

- In a reverse reaction, reactants are converted into products and products are converted into reactants. Over time, it will reach an equilibrium state where the concentration of the products and reactions are equal. This is measured by the equilibrium constant K .
- Before we reach K , we can calculate Q , which is the same (products/reactants). If we compare our K and Q values, it allows us to predict how the reaction will occur.

- When Q is:
 - Greater than K; equilibrium shifts left
 - Less than K; equilibrium shifts right
 - Equal to K; system is at equilibrium
 - *If you put the numbers on a number line it can help!!



Unit 8: Acid and Bases

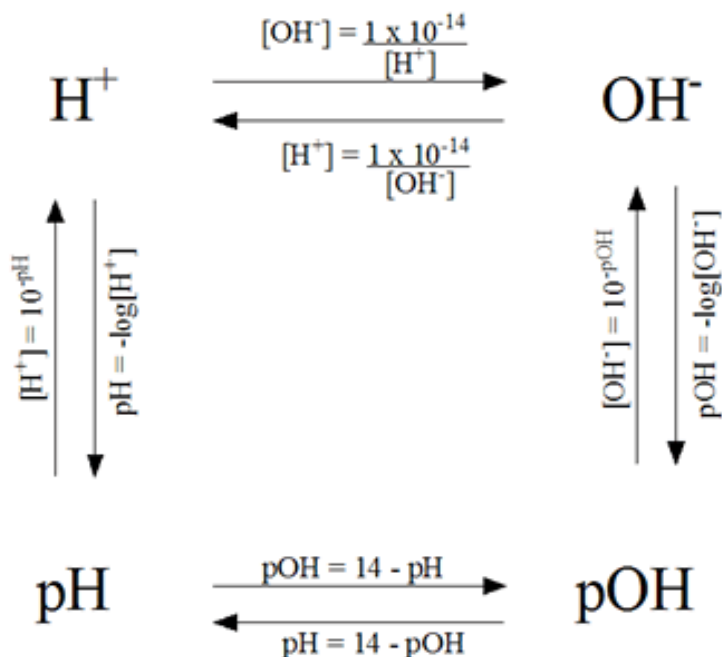
Strong Acids: HCl, HBr, HI, H₂SO₄, HClO₄, HClO₃, HNO₃. It's important that you memorize these.

- Strong Acids are able to completely dissociate into ions.

Strong Bases: LiOH, NaOH, KOH, RbOH, CsOH, Ca(OH)₂, Sr(OH)₂, Ba(OH)₂. You should memorize these also.

- Strong Bases are able to completely dissociate into ions.

A reaction with strong acids and strong bases create **salt and water.**



<http://www.sciencegeek.net/APchemistry/APtaters/pHcalculations.htm>

This chart represents the relationship between Acid and Base.

When solving a problem with acid or base, you use the ICE table like you did for normal equilibrium problems. However, it wouldn't be K , but rather, it would be K_a or K_b .

Buffers are solutions that can resist pH change. It can neutralize small amounts of acid or base, maintaining the pH.

Unit 9: Application of Thermochemistry

Entropy $\otimes S$: A driving force for a spontaneous process is an increase in the entropy of the universe. Entropy represents the disorder of a reaction.

- $S_{\text{solid}} < S_{\text{liquid}} < S_{\text{gas}}$
- Gas has the largest S because it has the most disorder between the gas molecules. On the other hand, solid has the smallest S because it has the least disorder between molecules in the solid.

Second Law of Thermodynamics: In any spontaneous process, there is always an increase in the entropy of the universe. Total energy of the universe is constant, but entropy is increasing.

$$\Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surrounding}}$$

Third Law of Thermodynamics: Entropy of a perfect crystal at 0 K is 0.

Standard Entropy Values (ΔS°): represents the increase in entropy that occurs when a substance is heated from 0 K to 298 K at 1 atm pressure.

$$\Delta S^\circ_{\text{reaction}} = \sum \Delta S^\circ_{\text{products}} - \sum \Delta S^\circ_{\text{reactants}}$$

Gibbs Free Energy ΔG : Total energy (entropy and enthalpy) that is available to do work. It determines the spontaneity of a reaction.

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

Standard Free Energy Change (ΔG°): The change in free energy that will occur if the reactants in that standard states are converted to the products in their standard states.

$$\Delta G^\circ_{\text{reaction}} = \sum \Delta G^\circ_{\text{products}} - \sum \Delta G^\circ_{\text{reactants}}$$

Here's a chart with relationship between enthalpy, entropy, and free energy:

ΔH	ΔS	$-T\Delta S$	ΔG	Spontaneity
+	-	+	+	Nonspontaneous
-	+	-	-	Spontaneous
-	-	+	+ or -	Low Temp: Spontaneous High Temp: Nonspontaneous
+	+	-	+ or -	Low Temp: Nonspontaneous High Temp: Spontaneous

<https://courses.lumenlearning.com/suny-introductory-chemistry/chapter/spontaneity-free-energy-and-temperature/>

There's also a relationship between Gibbs free energy (ΔG°) and K.

ΔG°	K
$\Delta G^\circ = 0$	K=1
$\Delta G^\circ < 0$	K > 1
$\Delta G^\circ > 0$	K < 1

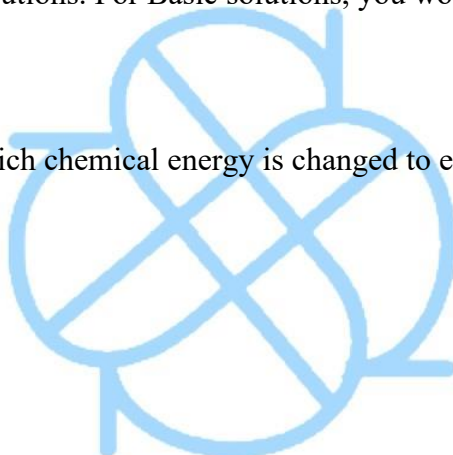
Electrochemistry is a study of a relationship between electricity and the chemical change.

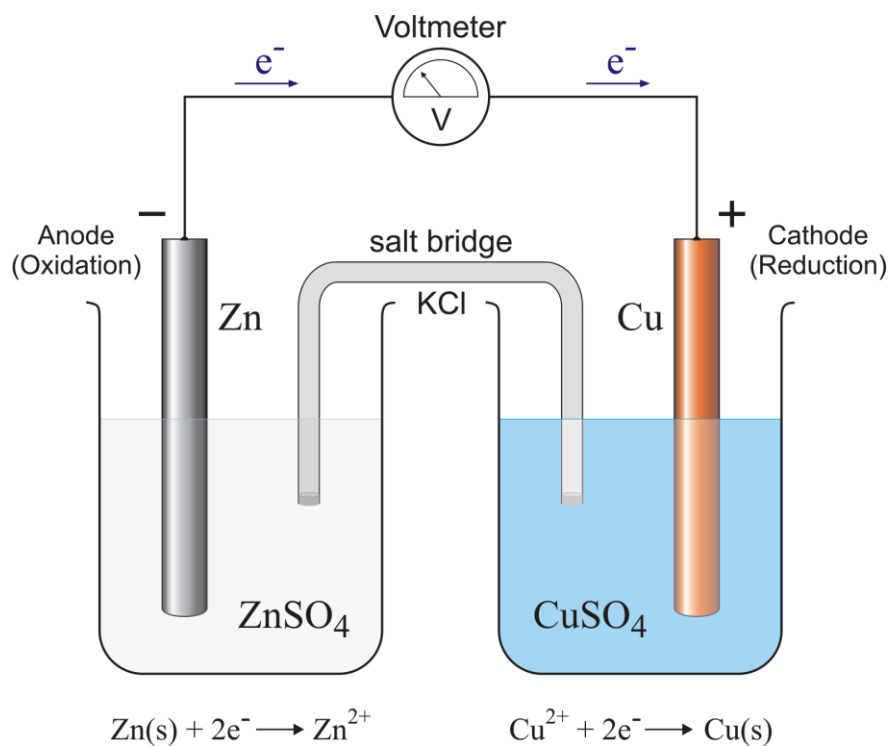
In past reactions, the electrons were already balanced on both sides. However, for this unit, we would have to balance the electrons and balance the equation overall. Here are the steps:

1. Split reaction into half - into oxidation or reduction
2. For each half reaction, balance all except H and O
3. Balance O using H_2O
4. Balance H using H^+
5. Balance charges using e^-
6. Multiply by an integer if the e^- doesn't cancel.
7. Add half-reactions and cancel identical species.
8. Check that elements and charges are balanced.

This process was for acidic solutions. For Basic solutions, you would just add OH^- to both sides and form water on both sides.

Galvanic Cell: a device in which chemical energy is changed to electrical energy. Here's the experiment:





<https://glossary.periodni.com/glossary.php?en=galvanic+cell>

There are also important formulas to memorize:

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$\Delta G^\circ = -nFE^\circ$$

n is balanced mole and F is a faraday's constant which equals 96485 C/mol e^-

Nernst Equation: $E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} \ln Q$

There's also a relationship between the Gibbs Free Energy, Equilibrium and the Galvanic cell:

ΔG°	K	E°_{cell}	
$\Delta G^\circ = 0$	$K=1$	0	Equally Favored
$\Delta G^\circ < 0$	$K > 1$	+	Products Favored
$\Delta G^\circ > 0$	$K < 1$	-	Reactants Favored