📋 AP Chemistry - Core Concept Master Cheat Sheet

01: Introduction to AP Chemistry

Use the KUDOS method for solving word problems.
K = Known
U = Unknown
D = Definition
O = Output
S = Substantiation

- **K (Known):** Use units to indentify information. Write information symbolically. Look for implied information. Write out chemical equations.
- **U (Unknown):** What is the problem looking for? Write information symbolically.
- **D (Definition):** Find equalities to convert. Choose & Re-arrange equations. Look for missing information in other places. If you cannot find enough information, re-evaluate your plan.
- **O (Output):** Plug in values to the equations (use constants as needed); Check unit cancellation & perform the calculation.
- **S (Substantiation):** Check validity of your answer. Check units, Check significant figures.

02: Fundamental Skills

The **metric system** uses prefixes to indicate multiples of 10
Metric Prefixes commonly used in chemistry

Prefix	Symbol	Multiple
Kilo	k	1000
Deci	d	0.1
Centi	c	0.01
Milli	m	0.001
Micro	μ	0.000001
Nano	n	0.000000001

The "base unit" is when there's no prefix.

Counting Significant Figures:
- If there is a decimal point anywhere in the number: Start with the first non-zero number and count all digits until the end.
- If there is not a decimal point in the number: Start with the first non-zero number and count until the last non-zero number.

Calculations with significant figures:
- Always complete calculations before rounding.
- Adding/subtracting: Answer has least number of decimal places as the problem.
- Mulitplying/dividing: Answer has least number of significant figures in problem.

Logarithms: Way of counting in multiples of the base

$$x = \log_b y \quad \text{and} \quad y = b^x$$

Using Dimensional Analysis:
1. Write your given information on the left side.
2. Write "= _____ (desired unit)" on the right side.
3. Find equalities that include both the desired unit and the given unit.
4. Arrange the equalities so that the given unit cancels.
5. Calculate answer (multiply across top and divide across bottom).

Multi-step Dimensional Analysis
- If there is no equality that contains both the given and the desired unit, you will need to use more than one equality.
- If you convert from a metric prefix to another metric prefix, use the base unit as a bridge in-between.

03: Matter, Energy & Changes

Matter			
Pure Substance		**Mixtures**	
Element	Compound	Homogeneous	Heterogeneous
Hydrogen	H2O	Tap water	Sand & Water

Energy	
Kinetic Energy (KE)	**Potential Energy (PE)**
Energy due to motion	Stored in chemical bonds

Physical Changes
- <u>**Do not**</u> create a new substance.
- All changes in state (between solids, liquids and gases) are physical changes.

Breaking, cutting, dissolving, drying, melting, freezing, etc.

Chemical Changes
- <u>Do</u> produce new substances.

Some signs of a chemical change are:
- production of a gas (bubbles)
- heat change (getting hot or cold)
- light
- change in color
- formation of a precipitate (forming an insoluble substance from two soluble substances.
- However, some of these signs could be present in physical changes as well.

Rusting, burning, reacting with water, reacting with acid, etc.

04: Atoms & Molecules

Sub-atomic Particles:

Particle	Location	Mass	Charge
Proton	Nucleus	1 amu = 1.67×10^{-27} kg	+1
Neutron	Nucleus	1 amu = 1.67×10^{-27} kg	0
Electron	Outside the nucleus	0.00055 amu = 9.10×10^{-31} kg	-1

Ions
- Atoms can gain or lose electrons to form **ions** (atoms with a charge.
- **Anion:** Atom with a negative charge.
- **Cation:** Atom with a positive charge.

Element symbols:

$$^A_Z X^C_\#$$ Where

- A = mass number (# of protons + # of neutrons) .
- Z = atomic number (# of protons) .
- C = charge (# of protons - # of electrons) .
- # = number of atoms.

Isotopes: Atoms of same element with different number of neutrons and different mass.
- Mass number refers only to a specific isotope

Calculating average atomic mass: (found on periodic table)
Atomic mass = Σ(fractional abundance)(mass of that isotope)

Atoms, elements and molecules
- **Atoms:** made of sub-atomic particles.
- **Elements:** made of the same type of atom (each has the same number of protons).
- **Molecules:** made of more than one type of atom (more than one element) chemically bonded together.

05: Writing Chemical Formulas

Type 1 Binary ionic: Contains two elements—one metal & one non-metal.
1. Write the symbol and charge of the first element.
2. Write the symbol and charge of the second element.
3. Balance the charges (to form a neutral compound) by using subscripts.

Type 1 or 2 with Multivalent Metals (metals that can have more than one charge).
1. The roman numberal indicates the charge of the cation metal.
2. Follow the rules for Type #1 or Type #2 as it applies.

Type 3 Binary Covalent: Contains two non-metals (which do not form charges when bonding together).
1. Do not worry about charges with this type.
2. Write the first element's symbol.
3. Write the second element's symbol.
4. Use the prefixes to determine subscripts ("mono" is not used on the first element).

Acids:
1. "Acid" indicates "H+" is the cation.
2. Choose the anion:
    a. "hydro___ic acid" – anion is single element (no oxygen).
    b. "___ic acid" – anion is "___ate" ion.
    c. "___ous acid" – anion is "___ite" ion.
3. Balance charges with subscripts.

06: Naming Chemicals

Type 1 Binary ionic: Contains two elements—one metal & one non-metal.
1. Write the name of the first element.
2. Write the name of the second element with "-ide" (subscripts do not matter in this type).

Type 2 Polyatomic Ionic: Contains at least one **polyatomic ion** (group of atoms that together have a charge).
1. Write the name of the metal or "ammonium" for NH_4
2. Write the name of the polyatomic anion (do not change the ending) or the single element with "-ide".
(subscripts within a polyatomic ion must match the name exactly. If there is parenthesis, the polyatomic ion is inside the parenthesis).

Type 1 or 2 with Multivalent Metals (metals that can have more than one charge)
Co, Cr, Cu, Fe, Hg, Pb, Sn
1. Name the cation and anion as for Type #1 or Type #2
2. The compound is neutral. Use the charge of the anion to determine the charge of the cation.
3. Write the charge of the cation in roman numerals inside parenthesis.

Type 3 Binary Covalent: Contains two non-metals (which do not form charges when bonding together).

1. Write the first element's name with the prefix indicating the # of molecules (mono- is not used with the first element).
2. Write the second element's name with the prefix indicating the # of molecules and "-ide".

Acids: (Compounds with "H+" cations are acids).
1. Look up the anion:
    d. No oxygen, a single element: "hydro___ic acid"
    e. "___ate" ion: "___ic acid"
    f. "___ite" ion: "___ous acid"

07: The Mole

Mole: SI unit for counting (abbreviation: mol).
- 1 mole of anything = 6.02×10^{23} pieces.
- The atomic mass found on the periodic table is the mass (in grams) for 1 mole of atoms of that element.
- At standard temperature and pressure (STP), 1 mole of any gas is 22.4 L **(Molar Volume of a gas)**.

Molar Mass (Molecular Mass, Formula Weight):
- By adding the atom masses for atoms in a molecule, the molar mass of the molecule can be found.
- Be sure to distribute subscripts outside the parenthesis to each atom inside.

Percent Composition:

$$\% \ composition = \frac{mass \ element}{mass \ whole} \times 100$$

If a chemical formula is given, use atomic masses and molar mass in % composition.

Empirical formula (lowest ratio of atoms in molecule):
1. If given percent's, assume they are grams. Change all grams to moles.
2. Divide all moles by the smallest to get the lowest ratio (multiply by a factor if needed to make whole numbers).
3. Write the formula with the ratio as subscripts.

Molecular Formula (actual ratio of atoms in molecule):
1. Find empirical formula, if not given to you.
2. Find the molar mass of the empirical formula.
3. Find the ratio of the molecular formula's molar mass (given to you) to the empirical formula's molar mass.
4. Multiple the empirical formula's subscripts by the ratio.

08: Chemical Reactions

Chemical Reaction: Bonds and atoms are rearranged to form new compounds.

Chemical Equation: Symbolizes the chemical reaction with chemical formulas.
- Reactants → Products
- States of matter are shown (s = solid, l = liquid, g = gas, aq = aqueous).
- Coefficients give mole ratio.
- A double arrow (⇆) indicates a reversible, an equilibrium reaction.

Types of reactions:
- **Composition:** More than one type of matter combine to form one type of matter.
- **Decomposition:** One type of matter decomposes into more than one type of matter.
- **Single replacement:** A single element changes place with an ion in a compound.
- **Double replacement:** Two ionic compounds switch ions.
- **Neutralization reaction:** Double replacement reaction with an acid and a base as the reactants.
- **Redox reaction:** Reduction-oxidation reaction.
- **Precipitation reaction:** A precipitate is formed.

Solubility rules for determining precipitates:

Anion	Forms insoluble compounds with
NO_3^-	No common ions
CH_3COO^-	Ag^+
Cl^-, Br^-, I^-	$Ag^+, Pb^{2+}, Hg_2^{2+}, Ti^+$
SO_4^{2-}	$Ag^+, Pb^{2+}, Ba^{2+}, Sr^{2+}, Ca^{2+}$
CrO_4^{2-}	$Ag^+, Pb^{2+}, Ba^{2+}, Sr^{2+}$
S^{2-}	All anions except NH4+, columns 1 & 2
OH^-	All anions except NH_4^+, column 1, Ba^{2+} & Sr^{2+}
CO_3^{2-}, PO_4^{3-}	All anions except NH_4^+, column 1 (except Li^+)

NH_4^+, Na^+ and K^+ are soluble with all common ions.

09: Balancing Equations

- The Law of Conservation of Mass/Matter requires that a chemical reaction be balanced.
- Coefficients balance atoms in a chemical reaction and indicate the number of compounds in a reaction.

Inspection Method (to balance the most simple reactions):
1. Make a list of the elements in the reaction.
2. Count the # of each type of atom on each side.
3. Add coefficients to balance the number of atoms.
4. Determine the total charge of each side and use coefficients to balance charge.
5. When elements and charge are balanced, place a "1" in any empty coefficient location.

Oxidation # Method (for simple redox problems):
1. Determine the oxidation numbers of each atom.
2. Determine the net change in charge. Use it to determine the ratio of atoms that would cancel out the change.
3. Use the ratio as coefficients for the simplest compounds containing those elements.
4. Finish balancing by the inspection method.

Half-Reaction Method (For the most complex redox reactions):
1. Use oxidation numbers.
2. Write two half-reactions, 1 for reduction and 1 for oxidation.
3. Balance all elements except H and O using inspection.
4. **For an acid redox reaction:** Balance the O's by adding H_2O to the side needing more O.
5. **For a base redox reaction:** Balance O by adding twice as many OH^- to the side needing more O.
6. **For an acid redox reaction:** Balance the H's by adding H+ to the side needing more H's. For a base redox reaction: Balance H's by adding H_2O to the side needing more H's.
7. Add electrons to balance the charges.
8. Multiply half-reactions by factors cancel out electrons.
9. Add the two half-reactions back together.
10. Cross out anything that appears the same on both sides.

10: Predicting Products

Net Ionic Reactions
1. Separate all aqueous, soluble, ionic compounds into ions. Only subscripts within polyatomic ions remain—all other subscripts are changed to coefficients.
2. Cross out all spectator ions.
3. Re-write the equation with remaining ions.
4. The AP Exam will always have "reactions"—all the ions will never cancel out.

Double replacement—precipitation:
Switch cations with anions of the reactants.
Write the new compounds, balancing charges with subscripts.

Double replacement—acid/base:
Remove the proton from the acid and give it to the base.

Decomposition:
1. Break the reactant into more than one product.
2. Carbonate compounds form CO_2 as a product. Ammonium salts form NH_3 as a product.

Composition:
1. Metals and metal oxides added to water make bases.
2. Nonmetals and nonmetal oxides added water make acids.

Complex Ion Formation:
1. Know the common complex ions on the AP exam: $Ag(CN)_2^{-1}$, $Fe(SCN)^{+2}$, $Cu(NH_4)_3^{+2}$, $Zn(NH_4)_3^{+2}$, $Ag(NH_3)_2^{+1}$, $Ni(NH_3)_6^{+2}$, $Al(OH)_4^{-1}$, $Zn(OH)_4^{-2}$

Combustion:
1. When hydrocarbons are added to O_2, they always produce CO_2 and H_2O.

Redox:
1. All single replacement reactions are redox reactions.
2. Other common oxidizers are NO_3^{-1}, MnO_4^{-1}, $Cr_2O_7^{-2}$, and H_2O_2.

11: Stoichiometry

Stoichiometry: Using the mole ratio in the balanced equation and information about one compound to find information about another in the reaction.

Equalities used during dimensional analysis for stoichiometry:
- **Mole ratio in balanced equation:** Use to convert between moles of different compounds in the balanced equation.
- **Molar mass:** Used to convert between grams and moles.
- **Concentration:** Used to convert between moles and liters of a solution.

$$Molarity = \frac{moles\ solute}{L\ solution}$$

- **Molar volume of a gas:** Used to convert between moles and liters of a gas at STP.

Limiting reactant: Reactant that stops the reaction by running out first.
- Once a reactant has run out, the reaction will stop.
- Do stoichiometry for each given reactant quantity to the same product each time. Choose the calculation that gives the smallest amount of product.
- The reactant that produced the smallest amount of product is the limiting reactant.

Percent yield: compares the actual yield to the theoretical yield.

$$\%\ yield = \frac{actual\ yield}{theoretical\ yield} \times 100$$

12: Atomic Structure

Electron cloud: Area outside nucleus where electrons are located.
Energy levels: Electron cloud is divided into energy levels for electrons.
Subshells: Energy levels of electrons are divided into subshells of equal energy orbitals.
Orbitals: Subdivision of subshell. Each orbital can hold 2 electrons.

4 types of subshells:

	Subshell	Begins in level	# of orbitals	# of electrons
↓ higher energy	s	1	1	2
	p	2	3	6
	d	3	5	10
	f	4	7	14

Aufbau Principle: Fill shells from lowest energy to highest.
Hund's Rule: electrons are placed in each equal-energy orbital before doubling up to produce the lowest energy atom.
Pauli Exclusion Principle: Two electrons occupying the same orbital must be opposite spins (angular momentum).

Use the periodic table as a guide (read left to right):

1s			
2s			2p
3s			3p
4s		3d	4p
5s		4d	5p
6s	4f	5d	6p
7s	5f	6d	7p

3 types of electron configuration notation:
Boxes & Arrows: O (8 electrons): 1s [↑↓] 2s [↑↓] 2p [↑↓] [↑] [↑]
Spectroscopic: Br (35 electrons):
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$
Noble Gas: Br (35 electrons): $[Ar]\ 4s^2\ 3d^{10}\ 4p^5$

13: Utilizing the Periodic Table

Periodic Table: Tool for organizing the elements.
Periods: Rows on the periodic table.
Groups: Columns on the periodic table.
Periodicity: Predictable patterns and trends on the periodic table.

General trends in the period table

Trend	→ a period	↓ a group
Atomic Mass	Increases	Increases
Atomic Radii	Decreases	Increases
Ionization energy	Increases	Decreases
Electron Affinity	Increases	Decreases
Electronegativity	Increases	Decreases

Radii when forming a cation: There are now more protons than electrons. The pull of the protons on each electron is greater. Cations have smaller radii than their parent atom.

Radii when forming an anion: There are fewer protons than electrons. The pull of the protons on each electron is less.
Anions have larger radii than their parent atom.

14: Chemical Bonding Theories

Bond type	Happens between	Electrons are
Ionic	Metal & non-metal	Transferred
Covalent	Non-metals	Shared
Polar Covalent	Non-metals	Shared unevenly
Metallic	Metals	pooled

Polar covalent bond
When nonmetals bond covalently with a large difference in electronegativity
- Absolute value of differences:
    - g.   0 – 0.4 = covalent
    - h.   0.5 – 1.4 = polar covalent
    - i.   1.5 – 4 = ionic

Sigma (σ) bond: First bond between two atoms formed from head on overlap of orbitals.
Pi (π) bond: 2nd bond between two atoms formed from overlap of parallel p orbitals.

Hybridization: Atomic orbitals hybridize into orbitals of the same energy for all sigma bonds of the atom.

# of sigma bonds	Hybridization	Formed from
2	sp	1 s and 1 p
3	sp^2	1 s and 2 p
4	sp^3	1 s and 3 p

Molecular Orbital Theory: Molecules form new orbitals when bonding, rather than using the overlap of atomic orbitals.
Order of filling molecular orbitals:
$1\sigma\ 1\sigma*\ 2\sigma\ 2\sigma*\ 1\pi_a\ 1\pi_b\ 3\sigma\ 1\pi_a*\ 1\pi_b*\ 3\sigma*$
Bonding Orbital: Molecular orbital that pulls the nuclei together.
Antibonding Orbital (*): Molecular orbital that pushes the nuclei apart.
Bond Order: determines the type of bond (none, single, double, triple) based on bonding and antibonding electrons.

$$Bond\ Order = \frac{bonding - antibonding}{2}$$

15: Lewis Structures & VSEPR Theory

Valence Shell: Electrons in the outermost shell that bond.
Octet Rule: Atoms are most stable with a full valence shell.

Arranging Atoms in Lewis Structures
1. With only 2 elements, arrange symmetrically.
2. "COOH" is a carboxylic acid (both O's bond to the C and the H goes on one of the O's).
3. Hydrogen and halogens cannot go in the middle.
4. Other atoms in the order they appear in the formula.
5. Hydrogen and halogen atoms go around the element they are written next to in the formula.

Lewis Structure: A 2D representation of a molecule and its bonds.
1. Arrange the atoms as above.
2. Determine the # of valence electrons for each atom.
3. Draw the valence electrons—do not double up where a bond is going to form between two atoms.
4. Count to see if all atoms have full valences
5. If two atoms adjacent to each other do not have full valences, move in an electron from each to form a double bond. Repeat for triple bond if necessary. Move hydrogens as needed to allow double/triple bonds.

Exceptions to the Octet Rule:
1. Hydrogen and Helium can only hold 2 electrons Boron and Beryllium can be full with 6 electrons.
2. Any element in period 3 or below can have more than 8 electrons

Valence Shell Electron Pair Repulsion Theory (VSEPR):
Bonds and lone pairs (electrons) repel and arrange themselves in 3D as far away from each other as possible.

16: Gases

Assumptions of the KMT
1. Gases are made of atoms or molecules.
2. Gas particles are in rapid, random, constant motion.
3. The temperature is proportional to the average kinetic energy.
4. Gas particles are not attracted nor repelled from each other.
5. All gas particle collisions are perfectly elastic (they leave with the same energy they collided with).
6. The volume of gas particles is so small compared to the space between them that the volume of the particle is insignificant.

Symbols for all gas Laws:
P = Pressure; V = Volume; n = moles; T = Temperature (in Kelvin: K = °C + 273); R = Gas constant (8.31 L kPa/mole K or   0.0821 L atm/mole K); "a" and "b" = correction factors for real gases

Combined Gas Law: $\dfrac{P_1 V_1}{n_1 T_1} = \dfrac{P_2 V_2}{n_2 T_2}$

Dalton's Law of Partial Pressure: $P_{total} = \sum P_{of\ each\ gas}$

Mole fraction: $\chi_A = \dfrac{mole_A}{mole_{total}}$

Partial Pressure and mole fraction: $P_A = \chi_A P_{total}$

Ideal Gas Law: $PV = nRT$

Ideal Gas Law with Molar Mass: $PV = \dfrac{m}{MM} RT$

Ideal Gas Law with Density: $P = D \dfrac{RT}{MM}$

Real Gas Law: $\left(P + \dfrac{n^2 a}{V^2}\right)(V - nb) = nRT$

17: Liquids & Solids

Intramolecular forces: chemical bonds within a molecule.
Intermolecular forces (IMF): physical attractions between separate molecules.
- **London Dispersion Forces:** all molecules, temporary ganging up of electrons, weakest IMF
- **Dipole-Dipole Forces:** all polar molecules, medium strength.
- **Ion-Dipole Forces:** between a polar molecule and an ion.
- **Hydrogen Bonding:** extreme dipole with H on an N, O or F with another molecule with N, O, or F; strongest IMF.

Amorphous solid particles are "trapped" in place before they can arrange themselves into a repeating pattern.
Three types of crystalline solids:
- **Atomic solids**
   Metallic solids—closest packing of metal atoms.
   Network solids—one giant molecule. Each atom is covalently bonded to surrounding atoms.
- **Molecular solids**—strong covalent bonds within the molecular, weaker physical attractions between them.
- **Ionic solids**—electrostatic attraction between ions. Ions are stacked to minimize like-charge repulsions.

Changes in state involve breaking or forming IMF's
Boiling/Condensation Point: Vapor pressure of liquid = atmospheric pressure. Liquid and gas at equilibrium
Melting/Freezing Point: Vapor pressure of solid = Vapor pressure of liquid. Solid and liquid at equilibrium
Melting: $\Delta H = m \times H_{fus}$

Evaporating: $\Delta H = m \times H_{vap}$

18: Solutions

Solution: Homogeneous mixture;
Solute: Substance being dissolved;
Solvent: Substance doing the dissolving
Factors affecting Solubility:
- **Pressure:** Gases: as Pressure increases, solubility increases.
- **Temperature:** Gases: higher temperature is lower solubility. Most solids: higher temperature is higher solubility.

Concentration Measurements:

% by mass: $\% \ mass = \dfrac{mass \ solute}{mass \ solution} \times 100$

Molarity (M): $Molarity = \dfrac{moles \ solute}{L \ solution}$

Molality (m): $Molality = \dfrac{moles \ solute}{kg \ solvent}$

Dilution equation: $M_1 V_1 = M_2 V_2$

Electrolyte: compounds dissociate into ions when dissolved in water. Allows the solution to conduct electricity.

Colligative Property: Property of a solution that depends on the concentration of the solute particles.
- **Van't Hoff i-factor (i):** Factor describing how many particles is actually in the solution compared to how many molecules were added.
- **Vapor Pressure** of a solution is always lower than the pure solvent.
- **Raoult's Law:** $P_A = \chi_A P_A^{\circ}$

Boiling point of a solution is always higher $\Delta T_b = iK_b m$

Freezing point of a solution is always lower than the pure solvent. $\Delta T_f = iK_f m$

19: Kinetics

Kinetics: Study of reaction rates
In order for a reaction to occur, the molecules must: Collide with the correct orientation and the **Activation Energy** (minimum energy needed for reaction)

Factors affecting rate
- **Surface area**—as surface area increases, rate increases.
- **Concentration**—as concentration increases, rate increases.
- **Temperature**—as temperature increases, rate increases.
- **Catalyst**—presence of a catalyst increases rate.

Elementary Step: Chemical equation showing reactants in one collision and the products formed.
Reaction Mechanism: Series of elementary steps that add up to the overall reaction.

- k = **rate law constant**. Is different for each reaction at each temperature.
- [A] = concentration of reactant
- $[A]_0$ = initial concentration of reactant
- t = time

Rate Laws:

Order	Differential Law	Integrated Law
0	Rate = k	$[A] = -kt + [A]_0$
1	Rate = k[A]	$\ln [A] = -kt + \ln [A]_0$
2	Rate = k[A]^2	$1 / [A] = kt + 1 / [A]_0$

Half life ($t_{1/2}$): Time for ½ of the original reactants to disappear. Use integrated rate law and use $[A] = \frac{1}{2} [A]_0$

Calculating E_A
A = Arrhenius constant (fraction of collisions with correct orientation); E_A = Activation energy (unit of J/mole); R = 8.31 J/K×mole; T = temperature (in Kelvin. °C + 273=K)

$$k = Ae^{-\frac{E_A}{RT}}$$

20: Equilibrium

Reversible Reaction: Reaction that goes in both directions.
Equilibrium: When the rate of the forward and reverse of a reversible process are equal.
Dynamic equilibrium: The number of reactants and products do not change, but the reaction continues to occur in both directions.

Writing equilibrium constant expressions
- Concentration of products over concentration of reactants.
- Do not include pure solids or pure liquids.
- Use the coefficients of the balanced equations as powers.

Reaction Quotient (Q): When concentrations at any time are plugged into the equilibrium constant expression.
- If Q = K, it's at equilibrium.
- If Q > K, reaction proceeds towards reactants.
- If Q < K, reaction proceeds towards products.

Solubility Product (K_{sp}): Equilibrium constant for a dissolution reaction. Written just as any K is.

ICE charts a technique for organizing information in an equilibrium problem
- Make a table with the reactants and products across top
- Place "ICE" down the left hand side, for Initial, Change and Equilibrium.
- Fill in any given information from the problem.
- Use the balanced equation's stoichiometric ratio to determine the "change" row.
- Use "equilibrium" values to plug in to the K expression.

Le Chatelier's Principle: A system at equilibrium will re-adjust to reach equilibrium again when disturbed.

21: Acids and Bases

- **Arrhenius acid:** Produces hydronium ion in water.
- **Arrehnius base:** Produces hydroxide ion in water.
- **Hydronium ion:** H_3O^{+1};
- **Hydroxide ion**: OH^{-1}

Strong acids	HCl, HBr, HI, HNO_3, $HClO_3$, $HClO_4$
Strong bases	$NaOH$, KOH, $Ca(OH)_2$, $Ba(OH)_2$, $Sr(OH)_2$

For **polyprotic acids**, each successive proton is weaker than the one before. (H_2SO_4 has a strong 1st hydrogen.)

Writing equilibrium constants: Ratio of concentrations of products to reactants with balanced equation coefficients as the powers. (Do not include pure solids or liquids.)

$$K_w = K_a \times K_b$$

$$K_w = [H_3O^{+1}][OH^{-1}] \quad \text{at 25°C, } K_w = 1.0 \times 10^{-14}$$

The larger the K_a or K_b, the stronger the acid or base

pH: Logarithmic scale of acidity.

$$pH = -\log[H_3O^{+1}] \qquad [H_3O^{+1}] = 10^{-pH}$$

Solving pH problems
- For strong acids, assume the $[H3O+1]$ = the concentration of the acid
- For weak acids, use an ICE chart and the Ka for weak acids (use approximations for tiny Ka's).

Salt from
- Weak acid + strong base = Basic
- Strong acid + weak base = Acidic

Buffer: Weak acid or base and its conjugate that resists changes in pH when acid or bases is added

$$pH = pKa + \log\left(\frac{[base]}{[acid]}\right) \text{ where } pKa = -\log[K_a]$$

22: Thermodynamics

Thermodynamics: Study of heat changes.
Energy: The ability to do work or supply heat.
Heat (q): Flow of energy from a hot toward a cooler object.
Enthalpy (H): Takes into account internal energy, pressure and volume. Same as heat for open-air situations

Work $\quad w = -P\Delta V$

w = work (in J); P = pressure (in atm); $\Delta V = V_2 - V_1$ (in L)

Specific Heat Capacity (C_p): Amount of energy that 1 gram can absorb before increasing in temperature.
C_p for water: 4.18 J or 1.00 cal

$$\Delta H = m \times C_p \times \Delta T \qquad m = \text{mass}; \Delta T = T_2 - T_1$$

Calorimetry: $\quad \Delta H_{system} = \Delta H_{surroundings}$
T_2 of both system and surroundings are the same
Enthalpy of formation (H_f): Energy change when a compound is formed from its elements.

$$\Delta Hrxn = \sum H_f prod - \sum H_f react$$

Entropy (S): Disorder or random-ness
Free Energy (G): Takes into account enthalpy, entropy and temperature to determine spontaneity

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

- ΔG = spontaneous at that temperature
+ ΔG = spontaneous in the opposite direction at that temp
$\Delta G = 0$ at equilibrium

$$\Delta G° = -RT \ln K \quad \text{and} \quad \Delta G = \Delta G° + RT \ln Q$$

$\Delta G°$ = free energy change at standard state (1 atm and 25°C)
R = 8.31 J/mole×K; T = temperature (in Kelvin); K = equilibrium constant; Q = Reaction Quotient

23: Electrochemistry

Electrochemistry: The study of the inter-change between electrical and chemical energy.
Voltaic cell (or Galvanic cell): Uses a redox reaction to produce electricity.
Electromotive force, EMF (or Cell Potential): Difference of potential energy of electrons from before and after the transfer.
Standard reduction potential: EMF if hydrogen is used as the other half-reaction (Hydrogen is defined as "0").
Calculating EMF from standard reduction potentials:
    EMF = cathode – anode
    + EMF = spontaneous
Stoichiometry & Electrochemistry:
1 amp (A) = 1 Coulomb/sec (C/s)
1 Faraday (F) = 1 mole of e^{-1}
1 Faraday (1 mole of e^{-1}) = 96475 Coulomb (C)
Free energy = maximum work that can be done by system

$$\Delta G = -n \times F \times EMF$$

ΔG = free energy (in J); n = # of moles of electrons transferred; F = 1 Faraday (1 mole of e^{-1}) = 96475 Coulomb (C); EMF = cell potential.

Nernst equation: $\quad EMF = EMF° - \frac{RT}{nF} \ln Q$

EMF = cell potential at current conditions; EMF° = cell potential at standard state (1 atm & 25°C); R = 8.31 J/mole×K; T = temperature (in Kelvin); n = moles electrons transferred; F = 1 Faraday; Q = reaction quotient $\quad \ln K = \frac{n \times F \times EMF°}{RT}$

24: Descriptive & Organic Chemistry

- **Descriptive Chemistry:** Knowledge of descriptive facts about chemistry.
- **Colorimetry:** Study of colored compounds and the relationship between light absorbed and concentration.
- **Organic chemistry:** Study of carbon-containing compounds.
- **Hydrocarbon:** Compound with hydrogen and carbon atoms.
- **Functional Group:** Group of atoms on an organic molecule that affect the properties of that molecule.
- **Isomers:** Molecules with the same chemical formula but different structure.

Colored soluble compounds	Colored insoluble compounds
Soluble copper salts are blue/green	AgCl is white
Fe salts are red/brown	Chromate precipitates are orange
Cobalt salts are blue	Dichromate precipitates are yellow
Complex ions are often colored	Hydroxide precipitates are white

The most common, simple **functional groups**:
- Haloalkane: Halogen replacing a hydrogen on the hydrocarbon.
- Alcohol: "-OH" replacing a hydrogen on the hydrocarbon
- Carboxylic acid: "-COOH" replacing a hydrogen on a hydrocarbon.

Carbons are counted in the longest chain, starting from the carbon closest to the functional group.
Prefixes are used to denote number of carbons:
    1 = meth-
    2 = eth-
    3 = prop-
    4 = but-
    5 = pent-
    6 = hex-
    7 = hept-
    8 = oct-
    9 = non-
    10 = dec-