

You gotta know these!

AP Chem fact sheet
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Flame Colors

Sodium - Yellow

Potassium - Lilac

Cesium - Blue

Copper (II) - Green

Calcium - Red

Cobalt - Blue

Aluminum - White sparkles

Rubidium - violet

Strontium - Crimson

Barium - Green

Magnesium - White

Polyatomic Ions

NH_4^+ Ammonium

Hg_2^{+2} Mercurous/Mercury (I)

H_3O^+ Hydronium

MnO_4^- Permanganate

CrO_4^{2-} Chromate

$\text{Cr}_2\text{O}_7^{2-}$ Dichromate

AsO_4^{3-} Arsenate

CO_3^{2-} Carbonate

HCO_3^- Hydrogen carbonate

CH_3COO^- Acetate

$\text{C}_6\text{H}_5\text{COO}^-$ Benzoate

$\text{C}_2\text{O}_4^{2-}$ Oxalate

HC_2O_4^- Hydrogen Oxalate

ClO_4^- Perchlorate

ClO_3^- Chlorate

ClO_2^- Chlorite

ClO^- Hypochlorite

SO_4^{2-} Sulfate

SO_3^{2-} Sulfite

$\text{S}_2\text{O}_3^{2-}$ Thiosulfate

HS^- Hydrogen sulfide

HSO_3^- Hydrogen sulfite

HSO_4^- Hydrogen sulfate

PO_4^{3-} Phosphate

HPO_4^{2-} Hydrogen phosphate

H_2PO_4^- Dihydrogen phosphate

NO_3^- Nitrate

NO_2^- Nitrite

CN^- Cyanide

CNO^- Cyanate

SCN^- Thiocyanate

O_2^{2-} Peroxide

OH^- Hydroxide

H^- Hydride

First-year Chem

- **Intensive properties:** do not depend on amount
- **Extensive properties:** depend on amount
- **Physical change:** changes appearance only
- **Chemical change:** chemical formula changes

Sig-figs:

1. Zeros in front ≠ significant
 2. Zeros between or at end = significant
 3. Multiplication → same as one with fewest SFs
 4. Addition → as many as fewest decimal places
- **Homogenous:** same or uniform
 - **Heterogeneous:** different

Important experiments/history

- **Dmitri Mendeleev:** first periodic table
- **John Dalton:** Atomic theory
- **J.J. Thompson:** discovered electrons (cathode ray tube), plum pudding model
- **Rutherford:** discovered nucleus, electron cloud (gold foil experiment)
- **Heisenberg:** uncertainty principle
- **Bohr:** Bohr model, electron shells
- **Einstein:** photoelectric effect, $E=mc^2$ (light has mass, wave-particle duality)
- **Gance:** discovered $-\log[\text{AP}] = 5$

Carbon names

- 1) Meth
- 2) Eth
- 3) Prop
- 4) But
- 5) Pent
- 6) Hex
- 7) Hept
- 8) Oct
- 9) Non
- 10) Dec

Alcohols: R-OH

Ester: R-O-C=O-R'

Alkanal: R-CHO

Solubility Rules

Soluble

- 1) Group I and NH_4^+
- 2) NO_3 , CH_3COO^- , ClO_3^-
- 1) Halogens (except F), SO_4^{2-}
 - a. except with Ag, Hg (I), and Pb
- 3) Strong acids and bases

Insoluble

- 1) CO_3^{2-} , O^{2-} , Silicates, PO_4^{3-}
 - a. except Gr. I and NH_4^+
- 2) S^{2-} except Gr I, II, and NH_4^+
- 3) OH^- except strong bases or NH_4^+

Oxidation Rules

1. Elemental form = 0
2. Monatomic ion = its charge
3. Nonmetals are usually negative #'s
 - H is +1 with nonmetals and -1 with metals
 - Fluorine is always -1,
 - Halogens in Oxyanions have positive oxidation numbers
 - Oxygen is -2 except in peroxide (O_2^{2-}), where it is -1
4. Sum of O# in a neutral compound is 0
 - In polyatomic ion, sum of O# is the charge

Strong acids

- HClO_4 (Perchloric acid)
 - HI (Hydroiodic acid)
 - HBr (Hydrobromic acid)
 - HCl (Hydrochloric acid)
 - H_2SO_4 (Sulfuric acid, 1st diss. only)
 - HNO_3 (Nitric acid)
 - H_3O^+ (Hydronium)
 - Sometimes*: HClO_3 , HBrO_3 , HBrO_4 , HIO_3 , and HIO_4
- * if it is on the AP, it is strong

Strong bases

- KOH (Potassium hydroxide)
- $\text{Ba}(\text{OH})_2$ (Barium hydroxide)
- CsOH (Caesium hydroxide)
- NaOH (Sodium hydroxide)
- $\text{Sr}(\text{OH})_2$ (Strontium hydroxide)
- $\text{Ca}(\text{OH})_2$ (Calcium hydroxide)
- LiOH (Lithium hydroxide)
- RbOH (Rubidium hydroxide)

Anonymous facts

- HF dissolves glass (SiO_2)
- Solutions with zinc are clear
- Alkalis form A_2O and H_2 (g) when reacting with cold water
- CO_2 contributes to the greenhouse effect
- CaCO_3 is chalk (INSOLUBLE!)
- Gold and copper are the only two non-white metals
- Mercury is a liquid at room temperature (only metal)
- Strontium is used for fireworks
- Natural gas contains hydrocarbons, mostly methane (CH_4)
- Bucky ball is C_{60} , has hexagon and pentagon faces

Aqueous Reactions and Stoichiometry (Ch 4)

- Electrolyte: Solutions contain ions (Ex: Na^+ , Cl^-)
- Nonelectrolyte: Solutions contain no ions (Ex: sugars)
- Electrolytes conduct electricity, nonelectrolytes do not
- Strong electrolytes dissolve fully (Ex: $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$)
- Weak electrolytes are in equilibrium (Ex: HCH_3COO)
- Types of reactions
 - Precipitation (Metathesis/Exchange)
 - Oxidation/Reduction
 - Acid/Base
 - Indicators show change in pH
 - Phenolphthalein turns pink with bases

Thermochemistry (Ch 5)

- Work: Energy to move something
- Heat: Energy to change temperature
- Energy, heat and work are related
- Work-Energy principle
- One joule = $1\text{kg}\cdot\text{m}^2/\text{s}^2 = 0.001\text{ kJ}$
- Kinetic energy = $\text{KE} = \frac{1}{2}mv^2$
- $E_{\text{el}} = kQ_1Q_2/d$; electrostatic potential energy
- $K = 8.99 \cdot 10^9 \text{ J}\cdot\text{m}/\text{C}^2$
- Work = $F \cdot d = -P\Delta V$; also in Joules
- 1 cal = 4.184 Joules = 0.001 Cal
- First law of thermodynamics: Energy is conserved
- $\Delta E = \Delta E_f - \Delta E_i$
- $\Delta E = q + w$; q = heat, w = work
- $\Delta E = \Delta H - P\Delta V$
- Endothermic absorbs heat
- Exothermic releases heat
- State functions depend on the present state only
- Ex: Energy, temperature, ΔH
- Non-state functions depend on how you got there
- Ex: Work, q_p
- When ΔP , $q_p = \Delta H$
- $\Delta H_{\text{rxn}} = H_{\text{products}} - H_{\text{reactants}}$
- Enthalpy is an extensive property
- Calorimetry: $Q = m \cdot C \cdot \Delta T$
- In calorimeter, $q_{\text{rxn}} = -q_{\text{solution}}$
- Bomb calorimetry: $q_{\text{rxn}} = -C_{\text{cal}} \cdot \Delta T$ (intensive)
- Hess's law: $\Delta H_{\text{rxn}} = \text{sum of the } \Delta H \text{ for each step}$
- $\Delta H^\circ_{\text{reaction}} = \sum n_p H_f^\circ(\text{products}) - \sum n_r H_f^\circ(\text{reactants})$
- ΔH_f° of elements and diatomic gases = 0!
- Standard conditions: "°" means at 1 atm and 298 K

Constants

- 1 amu = $1.66 \cdot 10^{-24} \text{ g}$
- $N_A = 6.022 \cdot 10^{23} / \text{mol}$
- $K_b = 1.38 \cdot 10^{-23} \text{ J/K}$
- $q_{\text{electron}} = 1.602 \cdot 10^{-19} \text{ C}$
- Faraday = 96,485 C/mol
- $m_{\text{electron}} = 9.11 \cdot 10^{-28} \text{ g}$
- $m_{p/n} = 1.67 \cdot 10^{-24} \text{ g}$
- $R_1 = 0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$
- $R_2 = 8.314 \text{ J}/\text{mol}\cdot\text{K}$
- $h_{\text{Planck's}} = 6.626 \cdot 10^{-34} \text{ J}\cdot\text{s}$
- $c_{\text{light}} = 3.00 \cdot 10^8 \text{ m/s}$
- $\text{AP}_{\text{score}} = 5$

Chemical Thermodynamics (Ch 19)

- *Spontaneous*: reaction can proceed without added energy.
- *Nonspontaneous*: rxn needs added energy to proceed.
- Reverse of a spontaneous reaction is nonspontaneous
- Entropy, S , and change in entropy, ΔS , is a system's disorder.
 - Its units are in Joules per mole Kelvin, or $\text{Jmol}^{-1}\text{K}^{-1}$
- A *reversible* process is one where you can reverse the change that just happened and the system and the surroundings are the exact same thing.
- An *irreversible* process is one where you can't reverse the change without a difference in either the surroundings or the system. Spontaneous reactions are always irreversible, and therefore being irreversible is favorable.
- *Isothermic*: process happens at constant temperature.
- $\Delta S_{\text{system}} = q_{\text{system}}/T$ (Constant temperature)
- The Second Law of Thermodynamics
 - $\Delta S_{\text{total}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}}$
- Molecules can move in 3 ways
 - *Translational*, which is a repositioning of a molecule; (Throwing a ball; note that ideal gases have only this type of movement)
 - *Vibrational*, which is where the molecule is like dancing, or moving internally
 - *Rotational*, where the molecule is spinning.
- Microstates = W = number of different positions a system can have
- $S = k_b \ln(W)$; $k_b = 1.38 \cdot 10^{-23} \text{ J/K}$
- $\Delta S = k_b \ln(W_f/W_i)$
- Ways to increase entropy:
 1. Increasing the temperature, which will increase the speed of the molecules
 2. Increasing the Volume, which provides more room for the molecules to spread out
 3. More molecules or moles
- $S_{\text{gas}} > S_{\text{liquid}} > S_{\text{solid}}$
- Entropy gains in reactions:
 1. Forming gases from liquids or solids
 2. Forming liquids from solids
 3. Dissolving a solid into solutions
- Third Law of Thermodynamics: At 0 K, $S = 0$
- $\Delta S^\circ_{\text{reaction}} = \sum n_p S^\circ(\text{products}) - \sum n_r S^\circ(\text{reactants})$
- Gibbs free energy: $G = H - TS$; $\Delta G = \Delta H - T\Delta S$
- $\Delta G^\circ_{\text{reaction}} = \sum n_p G_f^\circ(\text{products}) - \sum n_r G_f^\circ(\text{reactants})$
- If $\Delta G > 0$, nonspontaneous; $\Delta G < 0$, spontaneous
- $\Delta G = \Delta G^\circ + RT \ln(Q)$; Q is the reaction quotient
 - $RT \ln(Q)$ is an adjustment factor
- $K_{\text{eq}} = e^{(-\Delta G^\circ/(RT))}$
 - K will be greater than 0 when $\Delta G < 0$
 - Use 8.314 J/mol-K for these two and keep your units!

Ideal gases (Ch 10)

- High temperature
- Low pressure
- Kinetic-Molecular Theory
 1. Random, continuous motion
 2. Volume is negligible compared to container
 3. Negligible forces between molecules
 4. Collisions are elastic (no loss of KE)
 5. KE is proportional to temperature (K)
- STP
 - 1 atm (101,325 Pa, 760 torr, 760 mmHg)
 - 273 Kelvin (0°C)
 - 1 mol = 22.4 L
- Formulas:
 - $PV=nRT$ (ideal gas equation)
 - $u_{rms} = \sqrt{3RT/M_{molar}}$
 - $KE =$
 - $P_1V_1 = P_2V_2$ (Boyle's law)
 - $V_1T_2 = V_2T_1$ (Charles' law)
 - $n_1V_1 = n_2V_2$ (Avagadro's law)
 - $P_1V_1T_2 = P_2V_2T_1$ (Combined gas law)
 - $P_{total} = P_1 + P_2 + \dots$ (Dalton's partial pressures)
 - $P_1 = X_1P_{total}$ (mole fractions)
 - $R_e\sqrt{M_{molar(f)}} = R_f\sqrt{M_{molar(e)}}$ (rate of effusion)
 - $P = nRT/(V-nb) - n^2a/V^2$ (van der Waals)
 - Adjustments for occupying volume, collisions

Solutions (Ch 13)

- Solvent-solute attraction is called solvation
- When above process is in water, it is called hydration
- Crystallization is the reverse of solvation
- Saturated: equal solute for amount of solvent
- Unsaturated: excess solvent, can dissolve more
- Supersaturated: excess solute, solid on bottom
- Miscible: liquids mix in all proportions
- Immiscible: liquids do not dissolve completely
- Henry's law: Solubility of a gas is proportional to the pressure, $S_{gas} = kP_{gas}$
- Mass percentage: mass solute / mass solution * 100
- Parts per million (ppm): $m_{solute}/m_{solution} * 10^6$
- Parts per billion (ppb): $m_{solute}/m_{solution} * 10^9$
- Molarity (M) is moles solute / Liter solution
- Molality (m) is moles solute / kg solvent
- Colligative property: depends on concentration
- Ideal solutions obey Raoult's law, differences result from solvent-solvent and solute-solute forces
- i is the van Hoff's value (ions in solution)
- $\Delta T_b = iK_b m$ – boiling point elevation
- $\Delta T_f = iK_f m$ – boiling point depression
- This means that with a solution, there is a higher tendency to stay as a liquid
- Osmosis: movement of solvent through a semi-permeable membrane, from less concentrated to more concentrated
- Π (Osmotic pressure) = MRT
- Colloids: between solutions and mixtures
- Can be hydrophilic (like H₂O) or hydrophobic (dislike H₂O)
- Tyndall effect: scattering visible light with colloids

Intermolecular forces (Ch 11)

Crystalline solids

1. Covalent network
 - a. Atoms in a network of covalent bonds
 - b. Very hard
 - c. High melting point
 - d. Poor thermal and electrical conduction
2. Metallic
 - a. Metallic bonds
 - b. Soft to very hard
 - c. Low to very high melting point
 - d. Good thermal and electrical conduction
 - e. Malleable and ductile
3. Ionic
 - a. Electrostatic attractions (+/-)
 - b. Hard and brittle
 - c. High melting point
 - d. Poor thermal and electrical conduction
4. Molecular
 - a. Atoms or molecules
 - b. Fairly soft
 - c. Low to moderately high melting point
 - d. Poor thermal and electrical conduction

- Amorphous solid: random arrangement
- Crystal lattice: imaginary points where the unit cells are placed to obtain full form
- Unit cell: base unit that repeats over and over
- Primitive cubic: $1/8$ atom at 8 corners (1 atom/unit)
- Body-centered cubic: one atom at center and $1/8$ atom at 8 corners (2 atoms/unit)
- Face-centered cubic: $1/2$ atom at 6 faces and $1/8$ atom at 8 corners (3 atoms/unit)
- If an atom is on the edge, it counts as $1/4$
- Close packing is when they get close to maximize IMFs and minimize socializing
- There is single layer, cubic, and hexagonal

Intermolecular forces (Strongest to weakest)

1. Metallic bonds (except diamond is strongest)
 2. Covalent network
 3. Ion-ion
 4. Ion-dipole (happens in solutions)
 5. H-bonds (N, O and F only)
 6. Dipole-dipole
 7. London dispersion forces (all have these)
 - * see flowchart 11.2 on p452
- Things that increase strength of IMFs:
1. Polarity
 2. Polarizability (strength needed to remove electron from cloud)
 3. Size
 4. The shape of molecules affects IMF also
 - **Viscosity:** resistance to flow, in kg/m-s
 - **Surface tension:** tendency of a liquid to minimize surface area and maximize effort.

Equilibrium (Ch 15)

- Chemical equilibrium: forward and reverse processes occur at the same rate, neither direction is spontaneous
- Haber process: $\text{N}_2(\text{g}) + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3(\text{g})$
- Law of mass action: for $a\text{A} + b\text{B} \rightleftharpoons d\text{D} + e\text{E}$ has equilibrium expression of

$$K_c = \frac{[\text{D}]^d [\text{E}]^e}{[\text{A}]^a [\text{B}]^b}$$

- NO SOLIDS OR LIQUIDS! Only aqueous and gases!
- ICE the problem (Initial, Change, Equilibrium)
- If you want the pressure constant (K_P), substitute pressure (in atm) for molarity
- To convert from K_P to K_c , use $K_P = K_c(RT)^{\Delta n}$
- Same phase: homogenous equilibria
- Different phase: heterogenous equilibria
- Q is the reaction quotient. Take what you have at the beginning and plug into your K expression
- $K > 1$: favors products
- $K < 1$: favors reactants
- $Q = K$, at equilibrium
- $Q > K$, move to the left
- $Q < K$, move to the right

Le Châtelier's principle

If a system at equilibrium is disturbed, the equilibrium will shift to minimize the disturbing influence and consume excess

1. Increase temperature, rxn moves in the endothermic direction
2. Increase molarity of something, system goes the other way
3. Increase pressure shifts to decrease moles (Boyle's law)
4. Add volume shifts in the direction that yields the most moles
 - a. Your tire gets a hole in it. You take it to the shop and you fix it. Now, you're driving, and your tire explodes! How do you fix it? YOU CAN'T! Because Δn is 0.
 - b. If Δn is 0 for a system with a change in P or V , it is not fixable. Remember point 4(a)

Acid-Base Equilibria (Ch 16)

- Acids (H^+) turn litmus red
- Bases (OH^-) turn litmus blue
- Bronsted-Lowry acid: proton donor
- Bronsted-Lowry base: proton acceptor
- Lewis acid: electron-pair acceptor
- Lewis base: electron-pair donor
- Water is **amphiprotic** (acid or base)
- Conjugate acid: B-L base with H^+ added
- Conjugate base: B-L acid with H^+ subtracted
- Strong acids have weak conjugate bases, etc.
- $K_w = K_a K_b = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$
- $\text{pH} = -\log[\text{H}^+]$; $\text{pOH} = -\log[\text{OH}^-]$; $\text{pOH} + \text{pH} = 14$
- Weak acids and bases go into equilibrium
- For acids, use K_a ; for bases, use K_b (ref Ch 15)
- Still use ICE, but assume 5%. (Ex: $(0.015 - x) = 0.015$)
- To check 5% rule, take the $[\text{ion}]/[\text{WA}/\text{WB}] \times 100\%$
 - If this is less than 5%, you are good.
- When ions react with water, **hydrolysis** occurs
 - This changes pH; Spectators do not do this
- Strong acids and bases are one-way reactions.
 - IF these (Initial, Final)
Oxyacids (HIO , HClO_4 , H_2SO_4 , etc.)
- Structure is a nonmetal, X, bonded to (OH^-)
- These can be acids because the X-O bond is stronger than OH- bond and a proton leaves.
- Acidity increases as # Oxygens increase and electronegativity of X increases.

Additional Equilibria (Ch 17)

- Common-ion effect: Weak acid/base in solution with strong electrolyte that contains the same conjugate base/acid. (Ex: HCH_3COO and $\text{NH}_4\text{CH}_3\text{COO}$)
- This creates a buffer. The conjugate base/acid will try to keep the pH steady.
- Henderson-Hasselbach equation (it is a hassle to me)
 - $\text{pH} = -\log[K_a] + \log[\text{base/acid}]$
 - $\text{pOH} = -\log[K_b] + \log[\text{acid/base}]$
 - Only can use this BEFORE equivalence point in titrations
- At equivalence point of titrations, you have the conjugate base/acid remaining, which goes into another equilibrium. Use $K_w = K_a K_b$ for conjugates
- After equivalence point, there will be either OH^- for weak acid/strong base titrations, or H^+ for weak base/strong acid titrations. Take use the pH or pOH formula to calculate pH.
- Remember: you can neglect the remaining conjugate base/acid left after equivalence point because it is negligible.
- Also, you have to divide by total volume to get mol/L!

Steps for solving a problem

1. First thing you do: look at what they're giving you
2. Next, identify your substances (Strong acid, weak base, etc.)
3. Thirdly...



4. Convert to moles
5. You don't even know what they're asking yet!
6. ICE the problem
7. Do your math and get your answer.

K_{sp} and slightly solubles

- This time, you "IFE" the problem: Initial, Final, Equilibrium
 - If they give you the solubility, that is "X"
 - If they give you K_{sp} , they want you to find X. You've got to look at what they're giving you
 - Only the solid has initial. That's X
 - The solid has a final of 0, and no equilibrium value
 - The initial of the ions is 0, they have no final
 - The ions have an equilibrium of +X, +2X, etc. (varies by stoichiometry)
 - Remember: OMIT SOLIDS and liquids from K expression!
 - OH^- is usually insoluble. If they want pH, use $\text{pOH} = -\log[\text{OH}^-]$
 - If $Q > K_{sp}$, precipitate will form
 - If $Q < K_{sp}$, some precipitate will dissolve
 - If $Q = K_{sp}$, solution is at equilibrium
-
- Amphoteric oxides and hydroxides are slightly soluble in water but dissolve when an acid or base is added
 - Qualitative analysis: presence or absence
 - Quantitative analysis: how much is present

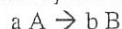
Kinetics (Ch 14)

(My colleagues and I predict this will be on the 2009 test. It is on there every other year, and it wasn't on there last year)

Things that affect the rate of reaction

- Orientation
- Surface area (for solids)
- 3. Temperature (increases speed for more intermolecular penetration)
- 4. Collisions
- 5. Concentrations/molarity
- 6. Catalyst (reduces activation energy)

Rate expressions



$$\text{Average rate} = (-1/a)\Delta[A]/\Delta t = (1/b)\Delta B/\Delta t$$

$$0 \text{ order rate} = -\Delta[A]/\Delta t = k[A]^0 = k; A = kt$$

$$1^{\text{st}} \text{ order rate} = -\Delta[A]/\Delta t = k[A]^1$$

$$\ln[A]_t = -kt + \ln[A]_0 \quad \text{or} \quad \ln[A_t/A_0] = -kt$$

$$t_{1/2 \text{ life}} = \ln(2)/k = 0.693/k$$

$$2^{\text{nd}} \text{ order rate} = -\Delta[A]/\Delta t = k[A]^2$$

$$1/[A]_t = kt + 1/[A]_0$$

$$t_{1/2 \text{ life}} = 1/(k[A]_0)$$

The fraction of molecules that have E_a (Activation energy) for a reaction is given by $f = e^{-E_a/(RT)}$; $R = 8.314 \text{ J/mol}\cdot\text{K}$ and you've got to keep your units!

Arrhenius equation, $k = A e^{-E_a/(RT)}$

k is the rate constant, and A is the frequency factor

$$\ln(k_2) - \ln(k_1) = -E_a/R(1/T_2 - 1/T_1)$$

The slope of the $\ln(k)-1/T$ curve = $-E_a/R$

- Elementary reactions occur in **one step**
- Two-step mechanisms involve two steps
 - The rate of the reaction depends on the slowest step in the reaction. (Rate-determining step)
 - A molecule created in one step and consumed in the next step is called an **intermediate**
- **Molecularity**: how many molecules take part in the reaction
 - Unimolecular: one molecule
 - Bimolecular: two molecules
 - Termolecular: three molecules
- Homogenous catalyst: same phase as reactants
- Heterogenous catalyst: different phase as reactants

Electrochem (Ch 20)

Hot tips:

- It is critical you know your oxidation rules
- You must identify your compounds
- What goes to what?
 - MnO_4^- goes to Mn^{+2}
 - CrO_4^{2-} or $\text{Cr}_2\text{O}_7^{2-}$ goes to Cr^{+3}
 - NO_3^- goes to NO_2

Steps to balance an oxidation-reduction reaction

1. Eliminate your spectators!
2. Separate into half reactions
3. Balance NON- Hydrogen or Oxygen atoms
4. Balance Oxygen with H_2O
5. Balance Hydrogen with H^+
 - a. If $\text{pH} > 7$ (basic solution), add as many OH^- as H^+ to both sides
6. Balance charge with electrons
7. Equalize $\frac{1}{2}$ reactions
8. Add the two reactions together
 - a. If you can **cancel the electrons**, you have done it correctly

Electrochemical cells

- Reduction happens at the cathode
- Oxidations happens at the anode
- e^- flow from anode to external circuit to cathode
- Voltaic cell: Uses spontaneous reactions to generate electricity (using its Gibbs Free Energy)
- Salt bridge allows anions to go to anode, cations to cathode
- $E^\circ_{\text{cell}} = E^\circ_{\text{red}}(\text{cathode}) - E^\circ_{\text{red}}(\text{anode})$
- When $E > 0$, spontaneous
- $\Delta G = -nFE$; $F = 96,485 \text{ C/mol}$
- $E = E^\circ - (RT/nf)\ln Q$; $Q = \text{reaction quotient}$
- Concentration cell is a voltaic cell where the same half reaction occurs in different concentrations
- 1 atm is standard for gases; 1 M for solutions
- **Cathodic protection**: Protecting a metal by putting it into contact with another metal that more readily undergoes oxidation
- When a battery is added, it is nonspontaneous
- **Fuel cells** are voltaic cells that use redox reactions where reactants are continually supplied to generate voltage
- Electrolysis: positive terminal to the anode, negative terminal to the cathode