

AP Chemistry

Thou Shalt Not Forget

Adapted from the original document by Dan Reid

Unit 1: Atomic structure and properties

1. When an electron is in a higher the energy level, it is farther away from the nucleus and therefore has less Coulombic attraction to the nucleus and is therefore easier to remove (...it has a lower 1st ionization energy.)
2. Moving across a row on the periodic table, the Z_{eff} increases, therefore the valence electrons are more attracted to the nucleus, therefore the atomic radius decreases and the ionization energy increases.
3. When reading a PES graph, the higher the peak, the more electrons there are in that sublevel, and a larger binding energy means that the electrons are closer to the nucleus.
4. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
5. When writing the electron configuration for a cation, remove the valence electrons first...the ones in the p-orbital and s-orbital...then you can remove d-orbital electrons if necessary. (note: if the last orbital filled using the Aufbau principle was a d orbital, the s orbital electrons are still the first to be removed).
6. Isotopes of an element have the same number of protons, but different numbers of neutrons.
7. Mass spectroscopy graphs measure atomic masses of isotopes.
8. Elements in the same group (vertical columns) have similar chemical and physical properties.
9. Metals are on the left side of the zig-zag line and nonmetals are on the right side of this line on the periodic table.
10. Cations (+) are smaller than their atoms since you are removing valence electrons that are farther from the nucleus; cations also have a higher Z_{eff} , so a greater positive force pulling on electrons. Anions (-) are larger than their atoms since adding extra electrons increases electron-electron repulsions.

Unit 2: Molecular and ionic compound structure and properties

1. Covalent bonds are formed between two nonmetals sharing electrons.
2. Ionic bonds are formed when a metal transfers electrons to a nonmetal and the opposite charges attract.
3. The greater the electronegativity difference between 2 atoms, the more polar the bond becomes.
4. Combustion reactions make CO_2 and H_2O .
5. Carbon makes a total of 4 bonds in a compound.
6. Bond angles: 4 domains = 109.5° 3 domains = 120° 2 domains = 180°

7. Hybrid orbitals: 4 domains = sp^3 3 domains = sp^2 2 domains = sp
8. Asymmetrical molecules = dipoles DO NOT cancel = polar molecule; symmetrical = dipoles cancel = nonpolar molecule
9. Single bond = sigma double bond = sigma + pi bond triple bond = sigma + 2 pi bonds
10. Lattice energy is the energy to break an ionic bond in a compound. Lattice energy increase as the ion's charge increases. Lattice energy decreases as the radii of the ions increase. (This can be deduced from Coulomb's Law.)
11. Formal charge involves comparing the # of valence electrons an atom has to the # of electrons around it in the Lewis structure. (Remember to "split" the bonded electrons evenly between the atoms.)
Formal charge = valence electrons – lone pair electrons -1/2 bonding electrons
12. Obey the octet rule first when drawing the Lewis Dot Structure then use formal charge if necessary. Extra electrons can go on the larger central atom, and if you have too few electrons, start making some double or triple bonds.
13. Metallic bonds are between metals, and they ALWAYS conduct electricity, and their hardness varies.
14. Interstitial alloys are made when a smaller atom fits into the gaps between the larger atoms of a metallic crystal. Substitutional alloys are made when the radii of the metals are similar in size and are substituted into the crystal lattice.

Unit 3: Intermolecular forces and properties

Intermolecular forces

1. IMF's from weakest to strongest: London Dispersion, dipole-dipole, hydrogen bonding, ion-dipole.
2. All molecules contain LD forces, and this force gets stronger as the molecule is larger...Larger electron cloud = more LD = more polarizable.
3. All polar molecules contain dipole-dipole forces, and this force gets stronger as the molecule is more polar.
4. H-bonds are between a NOF in one compound to a hydrogen that's already bonded to a NOF in another compound.
5. Boiling point and melting point increase as IMF's increase.
6. Vapor pressure and volatility decrease as IMF's increase.

Structures of solids

1. Molecular solids have low melting/boiling points, and they do not conduct electricity.
2. Ionic solids have high melting/boiling points, and don't conduct electricity as a solid, but DO conduct as a liquid or (aq).
3. SiO_2 (quartz) and diamonds are covalent network solids, and they have very high boiling/melting points.
4. When a molecular solid melts or boils, it is the IMF's between the molecules that break, not the covalent bonds.

Gases:

1. Gas mixtures are homogeneous b/c of the constant random motion of the particles.
2. Gases are compressible b/c of the large spaces between the particles.
3. Gas pressure is caused by collisions of particles with the walls of the container. More Collisions = More Pressure
4. P and V are inversely related...doubling the volume of a container will cut the pressure of the gas in half.
5. T and V are directly related...If you heat a balloon, it will expand.
6. T and P are directly related...If you heat a rigid container, the pressure of the gas will increase.
7. $PV=nRT$ Units: Temperature = Kelvin; Volume = Liters; Pressure = atm Use this gas constant $\rightarrow R=0.08206$
8. One mole of an ideal gas = 22.4 Liters ONLY at STP!!
9. Gas pressure and # of moles are directly related...if you double the mole of gas in a container, the pressure will double.
10. Molar Mass = dRT/P The "d" stands for density in units of g/L Use this gas constant $\rightarrow R=0.08206$
11. The more molar mass a gas has, the slower it moves at a given temperature.
12. Temperature = Average Kinetic Energy (Gases at the same temperature have the same average kinetic energy.)
13. When collecting a gas by water displacement: $P_{\text{total}} = P_{\text{dry gas}} + P_{\text{water vapor}}$
14. Real gases behave most like an ideal gas at high temperature and at low pressure. The more polar a gas is and the larger a gas is, the more it will **deviate** from ideal behavior. Consequently, small, nonpolar gases are the most ideal.

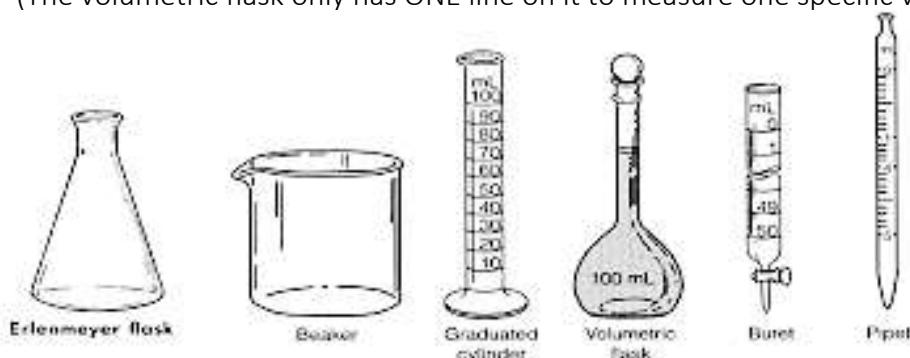
Solutions and mixtures

1. Compounds can be separated into elements by chemical changes, and mixtures can be separated by physical changes.
2. Filtering separates mixtures based on differences in particle size...the large particles are trapped on the filter paper while the soluble component goes through the filter paper and stays in the "filtrate".
3. Distillation separates mixtures based on differences in boiling point.
4. Chromatography separates mixtures based on differences in polarity.
5. In paper chromatography, the component that is most similar in polarity to the "mobile phase" moves up the farthest.
6. Density = mass/volume
7. The % composition by mass for a pure compound does not change.
8. $M_1V_1 = M_2V_2$ This is not on the formula sheet, but it is extremely useful for dilution calculations.

Laboratory:

1. When reading a volume of a liquid in a container, you can estimate by reading in between the graduated markings. That can give you one more sig. fig. in your volume. Read at the bottom of the meniscus. For example, in the picture you would read this as 36.5 and not 37 or 37.0.
2. Ranking measuring devices from least precise to most precise → beaker, graduated cylinder, volumetric flasks, burette

(The volumetric flask only has ONE line on it to measure one specific volume.)



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Unit 4: Chemical reactions

1. $\text{H}_2\text{O}_2\text{N}_2\text{Cl}_2\text{Br}_2\text{I}_2\text{F}_2$ -- the diatomic elements (“I Have No Bright Or Clever Friends” or just “Honclbrif”). When they are in a compound, their # of atoms can vary.
2. Empirical formula rhyme →

% to mass
mass to mole
divide by small
times until whole

...Get the simplest whole # ratio of the moles (or atoms) in the compound.

3. The molecular formula for a compound is a whole # multiple of the empirical formula ratio.
4. % yield = (experimental/theoretical) x 100%
5. % error = (experimental - theoretical)/theoretical x 100%
6. The amount of product for a reaction is determined by the limiting reactant.
7. Mass is conserved during both chemical and physical changes.

Unit 5: Kinetics

1. In order for a reaction to occur, particles must collide at the correct orientation & with a minimum energy to break bonds...(This minimum energy is called the activation energy...the height of the "hill".)
2. How to write a rate law for an elementary step... $2A + B \rightarrow C + D$ Rate = $k[A]^2[B]^1$
3. Rate constant (k) Units: 1st order = s^{-1} ; 2nd order = $M^{-1}s^{-1}$
4. Graphs: 1st order is linear for $\ln[A]$ vs time; 2nd order is linear for $1/[A]$ vs time ...
Absolute value of the slope = k
5. Ways to speed up a reaction: (1) Add a catalyst...lowers the activation energy (2) Increase reactant concentration...more collisions (3) Increase surface area...more collisions (4) Increase pressure of gases...increases the concentration of the gas, so there are more collisions (5) increase temperature...more collisions AND more of them have the minimum activation energy.
6. $\frac{1}{2}$ life for a 1st order process: $t_{1/2} = 0.693/k$
7. A 1st order reaction has a constant half-life regardless of the initial concentration. (Radioactive decay is a 1st order process.)
8. The taller the "hill" (or activation energy) the slower the reaction.
9. The slow step (rate-determining step) will dictate the speed of the reaction, and this step will determine the rate law.
10. Reaction Mechanisms: Intermediates are produced in one step and used up in a later step.
11. Reaction Mechanisms: Catalysts are used up in one step, and produced in a later step.

Unit 6: Thermochemistry

1. Exothermic reactions: $(-)\Delta H$; feels hot; heat is a product; temperature goes up...(endothermic is the opposite.)
2. $\Delta H_{rxn} = \text{Bonds broken} - \text{Bonds formed}$...(reactant bonds are broken; product bonds are formed)
3. Breaking bonds is endothermic. Forming bonds is exothermic.
4. $\Delta H_{rxn} = \Delta H_{products} - \Delta H_{reactants}$...Don't forget to multiply by the coefficients!!
5. If a reaction is exothermic, then the bonds formed in the products are stronger/more stable than the reactant bonds.
6. Doubling a reaction? ΔH will double. Reversing a reaction? The sign for ΔH changes. Adding reactions? Add the ΔH 's.

Unit 7: Equilibrium

1. $K_{eq} = \frac{[\text{products}]^x}{[\text{reactants}]^y}$... x and y represent the coefficients in the balanced chemical equation.
2. Only (aq) and (g) appear in an equilibrium expression. Use [] for Molarity and (P_{gas}) for atm.
3. A large K_{eq} means that there are more products at equilibrium. A small K_{eq} means there are more reactants at equilibrium.
4. Reversing a reaction? $1/K_{eq}$ Doubling a reaction? $(K_{eq})^2$ Adding reactions? Multiply the K's together
5. Le Chatelier's Principle: It's all about determining Q!! If $Q > K_{eq}$, then the reaction shifts to the left, towards the reactants.
6. Catalysts and inert gases DO NOT shift an equilibrium.
7. Changes in pressure (caused by changing the volume of a container) can shift an equilibrium ONLY IF the # of gas particles are different on each side...An increase in the pressure favors a shift in the equilibrium towards the side with LESS moles of gas. (Reminder: As $V \downarrow$, $P \uparrow$)
8. Molar solubility is the moles/L concentration of the particular species in a saturated solution at equilibrium
9. Solubility Equilibrium: 2 ions... $K_{sp} = x^2$; 3 ions... $K_{sp} = 4x^3$ "x" = Molar Solubility in units of moles/Liter. Some people use "s" as the variable instead of x (see topic 7.11)
10. The larger the "x" value, the more soluble the salt is.
11. If $Q > K_{sp}$, a precipitate forms.
12. Group I cations (Na^+ , K^+ , Li^+) NH_4^+ , and NO_3^- salts are always soluble in water. These are usually the spectator ions in a chemical reaction

Unit 8: Acids and bases

1. The pH of acids are less than 7, and bases are greater than 7. The pH of pure water is only 7 when the temp. is 25°C.
2. Acids donate $[H^+]$; bases accept $[H^+]$.
3. The hydronium ion is H_3O^+ . $[H^+]$ is a proton.
4. Strong acids: HNO_3 H_2SO_4 $HClO_4$ and HBr , HI , HCl ... "NO SO ClO 3, 4, 4 and BrCl"
5. Strong bases: Group 1 hydroxides Group 2 hydroxides *Some Group II hydroxides are only slightly soluble, but whatever dissolves can completely ionize.
6. $pH = -\log [H^+]$ $[H^+] = 10^{-pH}$
7. The stronger the acid, the weaker its conjugate base.
8. Acid-Base reactions favor the direction of the "strong side" to the "weak side"...If $K > 1$, then the reactants are stronger.
9. $[H^+] = \text{Square Root of } M_a K_a \dots$ (This shortcut only works if "x" is really small compared to M_a . Also, don't use this shortcut if you are given the pH of the solution and you are asked to solve for K_a because the pH can be used to find "x" in the ICE box.)
10. "x" in the ice box calculation is $[H^+]$ for a weak acid, and $[OH^-]$ for a weak base.

11. % Ionization of a weak acid = $[H^+] / M_a$
12. % ionization increases as the acid concentration decreases...adding more water will increase the amount of ionization.
13. If a salt contains a conjugate base of a weak acid, the salt is going to be slightly basic...CBOWA's are (-) ions.
14. If a salt contains a conjugate acid of a weak base, the salt is going to be slightly acidic...CAOWB's are (+) ions.
15. If a salt contains conjugates of strong acid/bases, the ion is neutral. Example -- KBr is a neutral salt (KOH + HBr)
16. A larger K_a value means a stronger acid. A larger K_b means a stronger base.
17. Relative strengths of acids: (a) Smaller cations are more acidic. (b) More (+) charge on the cation makes it more acidic. (c) More oxygens (or more electronegative atoms) on an anion makes it more acidic since the proton is "more ionizable".

Titrations and buffers

1. Buffers are created by a weak acid + CB (salt) or by a weak base + CA (salt).
2. $[H^+] = M_a K_a / [salt]$...You can use # of moles instead of molarity in this formula.
3. Adding a common ion to a weak acid (or base) decreases the % ionization, and therefore the pH gets closer to 7.
4. $M_a V_a = M_b V_b$...This is only true at the equivalence point.
5. Titrations: Weak acid + Strong Base has a pH at the equivalence point that's above 7. Weak Base + Strong Acid has a pH at the equivalence point that's below 7. Strong Acid + Strong Base has a pH = 7 at the equivalence point.
6. $pH = pK_a$ at the $\frac{1}{2}$ equivalence point for a "weak + strong" titration. Also, when $pH = pK_a$, then $[HA] = [A^-]$
7. More buffer capacity = more moles of weak acid & CB (or weak base and CA).

Unit 9: Applications of thermodynamics

1. Thermodynamically favorable (spontaneous) reactions have a $(-)\Delta G$.
2. Reactions with $(-)\Delta H$ and $(+)\Delta S$ are ALWAYS thermodynamically favorable... "enthalpy driven & entropy driven"
3. Reactions that increase the # of moles of gas have a $(+)\Delta S$.
4. If ΔG is $(-)$, then $K_{eq} > 1$.
5. ΔH and ΔS are usually NOT given in the same units!! When using $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$, make sure they match units.
6. $\Delta G = 0$ at equilibrium.
7. When using $\Delta G^\circ = -RT \ln K$, the value for R is 8.314 J/mol K so the answer for ΔG will be in the units of Joules.
8. Sometimes a reaction with a $(-)\Delta G$ does not proceed at a measurable rate. They are said to be under "kinetic control." High activation energy is a common reason for a process to be under kinetic control.

Electrochemistry:

1. Oxidation #'s: H = +1 (except in a hydride when it is -1) O = -2 (except in a peroxide when it is -1).
2. LEO goes GER ... Oxidation always occurs at the anode in both a battery and an electrolytic cell.
3. Electrons in a battery flow from anode (-) to cathode (+).
4. Salt bridge: Cations flow to the cathode, and the anions flow to the anode.
5. While a battery is discharged, the cathode gains mass and the anode loses mass.
6. If you reverse a reaction, the sign of E°_{cell} changes, but if you double a reaction, E°_{cell} DOES NOT change!!
7. $E^\circ_{\text{cell}} = E^\circ_{\text{Red (GER)}} - E^\circ_{\text{Red (LEO)}}$ (The other way to calculate $E^\circ_{\text{cell}} = E^\circ_{\text{Reduction}} + E^\circ_{\text{Oxidation}}$...but that involves reversing one of the reactions and changing the sign for E°_{Red})
8. The half-reaction with a more (+) E°_{Red} is the reaction that takes place at the cathode...GER.
9. When adding the two half reactions together, the electrons MUST cancel out.
10. $\Delta G^\circ = -nFE^\circ$ If ΔG° is (-), then E°_{cell} is (+). Reminder: n = # of electrons transferred
11. If Q increases, then the voltage (E°_{cell}) of the battery goes down.
12. Electroplating/Electrolysis Calculation: grams = (Molar Mass of the metal)(amps)(seconds)/(n)(F) ... $g = (\text{MM})(I)(t)/nF$