CHEMISTRY 338

Semester 2 Study Guide – Version 1.10

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I. Introduction to Chemistry

Goggles	Goggles are necessary in order to protect the eyes and ocular region from hazardous chemicals. Goggles should be worn at all times when doing a lab or when one is near a lab in which hazardous chemicals are used. Even boiling water may require goggles. Goggles are usually made of transparent/translucent rubber or glass to offer a nice viewing angle while still protecting the eyes.
Accident Reporting	There are many kinds of accidents that may happen in the lab. These include spilling of chemicals, breaking of lab equipment, or accidental reaction between elements. When an accident occurs, the teacher or leader must be notified immediately. Then, if something is spilled on someone, be sure to follow directions regarding that problem.
Bunsen Burner lighting	Most of the time, this requires goggles to be worn. 1. Turn on the gas 2. Light burner with lighter 3. Turn controls to regulate flow
Physical and Chemical changes	**Physical changes** are those changes that do not result in the production of a new substance. If you melt a block of ice, you still have H_2O at the end of the change. Changes of state, like melting or condensing, are physical changes. **Chemical changes** are changes that result in the production of another substance.
Sig Figs	These digits in a number represent the degree of accuracy of that number. Every non-zero digit in the number is significant. **Rules for Determining Sig Figs:** 1. All non-zero digits are significant. 2. Zeros to the left of non-zero digits are never significant For example, the number 007 has one sig fig, and .000045 has two sig figs. 3. Zeros between non-zero digits are always significant. For example, .10005 has five sig figs. 4. Zeros to the right of non-zero digits are significant only if a decimal point is shown. For example, 600 has one sig fig, but 600. has three sig figs.
Sig Figs in arithmetic	When adding or subtracting numbers, round the answer off to the same column as the **least** precise measurement used in the calculation. For example, $1.2 + 3.456 \approx 4.6$ because 1.2 is the least precise measurement in the calculation. When multiplying or dividing, the answer must be rounded off to the same number of **sig figs** as the least accurate measurement used in the calculation. For example, $1.0 \cdot 2.88991010 \approx 2.9$ because 1.0 has 2 sig figs, while 2.88991010 has 9 sig figs. Therefore, round it to 2 sig figs.
Scientific notation	A number in scientific notation looks like this: $\square.\square\square \times 10^{\square}$ The number before the decimal point must be 1, 2, 3 … 9. It cannot be zero or greater than or equal to 10.

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Accuracy and Precision	Accuracy is how well the results match the accepted value. Precision is how well the results match each other.
Types of Observations	Qualitative: Looking at attributes Quantitative: Looking at the number of things
Metric Conversions	1 inch = 2.54 cm 1 gram = 0.0353 oz
Measurement/Units	Mass: g (gram) Length: m (meter) Volume: cm^3 (cubic centimeter), L (liter) Temperature: °C, °F, °K
Celsius/Kelvin conversion	°K = °C + 273.15°
Solubility Curve	This is how much of a substance can dissolve in water (verify?) at a given temperature. The temperature is the X axis; amount of substance is Y axis.
Density Curve	This is a relationship between mass and volume at a given temperature. Y axis is density, X axis is temperature.
Density Formula	Density = Mass / Volume
Density of water	1.00 g/cm^3

II. Atomic Theory and Structure

Dalton's atomic theory	1. All matter is made of atoms. Atoms are indivisible and indestructible. 2. All atoms of a given element are identical in mass and properties. 3. Different atoms are different elements. 4. Compounds are formed by a combination of two or more different kinds of atoms. 5. A chemical reaction is a rearrangement of atoms and form new compounds.
Modern atomic theory	Modern atomic theory is a little more than Dalton's atomic theory. Now, we know that there are different kinds of atoms (differing by their masses) within elements that are known as **isotopes**. In addition, atoms can be destroyed via nuclear reactions but not by chemical reactions.
Experiments	**Cathode Ray Tube**: The Cathode Ray Tube consisted of a glass cylinder an anode and a cathode, a power supply and two silver conducting wires. The cathode was the negatively charged side of the tube. The anode was the positively charged side. When some air was pumped out, the electricity passed through the tube. This proved that electrons exist. **Gold Foil**: When a thin sheet of gold foil was bombarded with alpha particles, most of them penetrated the gold foil, others were deflected slightly, while some of them even bounced back the way they came from. This proved that atoms are mostly empty space.
Atomic number	This is the number of protons an atom has.
Mass number	This is the number of protons plus the number of neutrons an atom has.
Isotopes	The notation of an isotope looks like this: ${}^{14}_{6}C$, which is Carbon-14. The top number, 14, symbolizes the mass number of the isotope, while the bottom number, 6, symbolizes the atomic number of the isotope.

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Average Atomic Mass	It is the average of the masses of all the isotopes times the relative percent abundance. (This is not clear, so it is explained later.) The relative percent abundance is the abundance of an isotope in nature compared to the abundance of the element in nature. For example, 98.9% of all the carbon in the world is Carbon-12, so Carbon-12's relative percent abundance is 98.9%. We will now calculate the average atomic mass of an imaginary element. Let us say that element *X* has 2 isotopes: *X*-5 and *X*-6. *X*-5's percent abundance is 80%, and *X*-6's percent abundance is 20%. The average atomic mass would be 5 (mass of *X*-5) times 0.8 (abundance of *X*-5) plus 6 (mass of *X*-6) times 0.2 (abundance of *X*-6).

III. Nuclear Chemistry

Types of nuclear reaction	There are five types of nuclear reaction: 1. Alpha emission: Release of ${}^{4}_{2}\alpha$ from the nucleus. The atom loses 2 protons and 2 neutrons. 2. Beta emission: Emission of an electron (beta particle) from the nucleus. The atom gains a proton and loses a neutron. 3. Positron emission: Emission of a positron from the nucleus. The atom gains a neutron but loses a proton. 4. Electron capture: The atom captures an electron, gaining a neutron but losing a proton. Same as positron emission. 5. Gamma emission: Occurs when the nucleus is excited. It does not do anything to the mass of the atom. i.e. Alpha (α) emission: ${}^{238}_{92}\text{U} \rightarrow {}^{234}_{90}\text{Th} + {}^{4}_{2}\text{He}$. Beta (β) decay: ${}^{234}_{90}\text{Th} \rightarrow {}^{0}_{-1}\beta + {}^{234}_{91}\text{Pa}$.
When emissions occur	Alpha emission: Number of protons is greater than 83. Beta emission: Number of neutrons divided by the number of protons is too big. Positron emission/Electron capture: Number of neutrons divided by the number of protons is too small. Gamma emission: when the nucleus is excited.
Balancing nuclear reactions	Example: ${}^{238}_{92}\text{U} \rightarrow {}^{234}_{90}\text{Th} + {}^{4}_{2}\text{He}$. Notice that the sum of the mass numbers on both sides are the same ($238 = 234 + 4$) and the sum of the atomic numbers on both sides are also the same ($92 = 90 + 2$). This must be true, or the nuclear reaction is not balanced.
Half life	This is the time it takes for a radioactive isotope to lose half its mass. i.e. If isotope X has a half-life of 1 day, and our sample has a mass of 500 g, in 1 day, there will be 250 g left; in 2 days, there will be 125 g left. Equation: Mass after *n* half-lives $= \text{initial mass} \cdot \left(\frac{1}{2}\right)^{n}$

IV. Periodicity

Names of families	The first family is called the Alkaline Metals. The second family is called the Alkaline Earth Metals. The 7^{th} family (second from the right) is called the Halogens. The 8^{th} family (rightmost one) is called the Noble Gases.
Trends on the periodic table	From left to right: ionization energy increases, electronegativity increases, and atomic

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	radius decreases. Also, melting/boiling points decrease. From top to bottom: Atomic radius increases, electronegativity decreases, and ionization energy decreases. Also, melting/boiling points increase. Also, metallic character increases when you go from right to left and top to bottom.
Atomic orbitals	An orbital is a region of space in which you can expect to find electrons of specific energy. Each orbital 'contains' two electrons. In the first energy level, we find one 's' orbital. In the second energy level, we find one 's' orbital and three 'p' orbitals. In the third energy level, we find 1 's', 3 'p', and 5 'd' orbitals. In the fourth energy level, we find 'f' orbitals.
Electron configurations	$1s^2 2s^1$ The 1 before the first *s* and the 2 between the *s*'s represent the energy level. Therefore, the first *s* is in the first energy level, and the second *s* is in the second energy level. The superscript 2 and 1 represent the number of electrons found in the orbital, which is in this case 1*s* and 2*s*. Each orbital has 2 spaces for electrons. One of these spaces has positive spin, while the other one has negative spin. If an atom has more than two electrons, the electrons begin filling orbitals in the next subshell with one electron each until all the orbitals in the subshell have one electron. The electrons that are left then go back and fill each orbital in the subshell with a second electron with opposite spin. They follow this order because it takes less energy to add an electron to an empty orbital than to complete a pair of electrons in an orbital. The electrons fill all the subshells in a shell, then go on to the next shell. The *d* orbitals require more energy to fill, so electrons actually fill in 4*s* before they fill in 3*d*. Also, the *f* orbitals require even more energy to fill, so 6*s* is filled before 4*f* is filled.
Exceptions to configurations	Cr is $[\mathrm{Ar}]4s^1 3d^5$. Cu is $[\mathrm{Ar}]4s^1 3d^{10}$. Mo is $[\mathrm{Kr}]5s^1 4d^5$. Ag is $[\mathrm{Kr}]5s^1 4d^{10}$. Au is $[\mathrm{Xe}]6s^1 4f^{14} 5d^{10}$.
Orbital box diagrams	Go to http://wine1.sb.fsu.edu/chm1045/notes/Struct/EConfig/Struct08.htm. This also has a little bit on electron configurations.
Pauli exclusion principle	No more than 2 electrons can occupy an orbital. When two electrons occupy the same orbital, they have different spins.
Hund's rule	Try to maximize the number of unpaired electrons. I.e. when there are 3 electrons in a set of 3 *p* orbitals, put each electron in a different orbital.

V. Bonding

Properties	Ionic compounds: Formula based, very high melting and boiling points, solids at room temperature, soluble in water Covalent compounds: Molecule based, liquids and gases at room temperature, weak forces between molecules, low melting points, does not conduct electricity
Polarity	A molecule is **polar** if its polar bonds (if it has any) do not cancel each other out. A bond is polar if the difference in electronegativity (look on the periodic table to find out electronegativity) is 0.4 or greater. 'Canceling out' occurs when there are no unshared electron pairs. Let us look at two examples: H_2O and H_2Be (the second may not exist, but it is a good example ☺) H_2O has 2 shared and 2 unshared electron pairs, like shown:

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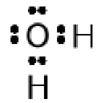
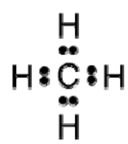
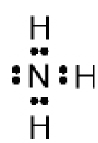
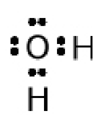
	 Since it has unshared electron pairs, and the O-H bond is polar, the molecule is polar. H_2Be looks like **H : Be : H**. Since it has no unshared electron pairs, even though the Be-H bond is polar, they cancel out and the molecule is not polar. In a polar bond, the electrons are closer to the atom that is more electronegative. The more electronegative atom is designated δ^-, or delta-negative, and the less electronegative atom is designated δ^+, or delta-positive.
Predictions due to Polarity	Polar substances: hydrophilic (dissolves in water). Things with oxygen or nitrogen in them are usually polar. Non-polar substances: hydrophobic, doesn't mix with polar substances. Hydrocarbons are always non-polar.
VSEPR theory	Valence Shell Electron Pair Repulsion. The whole concept revolves around the idea that the electrons in a molecule repel each other and will try and get as far away from each other as possible. Therefore, in models of the configuration, some angles are:   It is a central Carbon with 4 Hydrogens attached to it, so its shape is **tetrahedral** and its bond angle is 109.5°.  This is a central nitrogen with 3 hydrogens attached to it. It has 3 shared and 1 unshared electron pairs. Its shape is **trigonal pyramidal** (or just pyramidal). Its angle is 109.5° (actually it is 107°, but we said it was 109.5° in class).  This is a **bent** shape because it has 2 shared and 2 unshared electron pairs. Its degree angle is 105°.  This is a **linear** shape because it is straight. Its angle is 180°. Another shape (not shown here) is **trigonal planar**. It is a central atom with three shared pairs of electrons and no unshared pairs. Its angle is 120°.
Neutrally Charged Atoms to Ions	This occurs when an ionic molecule breaks down or dissolves in water or another substance. For example, when salt (NaCl) dissolves in water, it turns to Na^+ and Cl^-.

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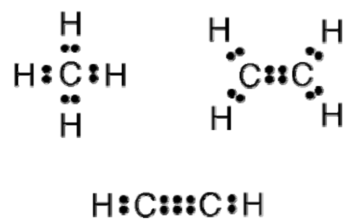
Electron dot structures	It shows the valence electrons of an atom. How to find it: 1. Count the valence electrons. Add them together; this is how many electrons you need to have in the final structure. 2. Assemble the bonding framework (draw the atoms by each other in the way you think they are arranged) 3. Place the electrons on the structure, making sure you follow the octet rule and its exceptions. Example: 

VI. Formula Writing

IUPAC names	This is a name system developed by the International Union of Pure and Applied Chemistry (IUPAC). It works by finding the longest chain of carbons, and then looking at what groups of atoms are attached to those chains. To see some examples of this, look at http://dl.clackamas.cc.or.us/ch106-01/iupac.htm.
Ionic compounds	In these, the charges of the anions (negative ions) and the cations (positive ions) must add to be zero. So, if the cation has a charge of +2 and the anion has a charge of -1, there needs to be 2 of the anion to make the charges add to zero. To write a ionic compound, first write the name of the cation. Nothing needs to be changed in it. Then, write the name of the anion like this: If the anion is an element, write the element with an ending of –ide. Therefore, chlorine would turn into chloride. If the anion is a compound, just write the name of the compound down. Therefore, NaCl will be sodium chloride, and K_2S will be potassium sulfide.
Binary covalent compounds	These are made up of two elements. Covalent bonds occur when two non-metals bond with each other. The subscript in the formula (i.e. 2 in H_2) turns into a prefix in the name. The prefixes are as follows: 1 – mono 2 – di 3 – tri 4 – tetra 5 – penta 6 – hexa 7 – hepta 8 – octa 9 – nono 10 – deca Therefore, SO_3 will be sulfur trioxide, and N_2O_5 will be dinitrogen pentoxide.
Monoatomic/Polyatomic ions	Monoatomic ions are either cations or anions that are made up of only one element. These include Na^+, Cl^-, and Al^{3+}. Polyatomic ions are cations or anions that are made up of more than one element. For example, SO_4^{2-} and NH_4^+ are polyatomic ions.
Percent Composition	Percent by mass of each element. Found by finding molar mass of compound, then the mass of the element in the compound per mole of compound, and dividing the second one by the first one. A less confusing explanation: Say you have some water and are trying to find the percent composition of oxygen. A mole of water's mass is 18.0152 g (found by adding molar masses of each element in water). There is one mole of oxygen in each mole of water, so we find the molar mass of oxygen, which is equal to 15.9994 g. Then, we divide:

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	$\frac{\text{Molar Mass of Element}}{\text{Molar Mass of Compound}} = \frac{15.9994}{18.0152} \approx 0.8881 = 88.81\%$.
Empirical Formulas	This is the lowest integral ratio possible for the elements in a compound. For example, the empirical formula for hydrogen peroxide (which is really H_2O_2) is HO. When the number of grams of each element in a compound is given, it is calculated as follows: 1. Find the number of moles of each element, given the number of grams of each. 2. Determine the mole ratio by dividing each number found in #1 by the lowest number. 3. Approximate the subscripts of each element using the mole ratio. (This is approximate, so if you get 1.98:1, you can approximate it as 2:1).
Molecular Formulas	This is found by calculating the mass of the compound from the empirical formula. Then, divide the total mass of the compound given in the first place by this empirical mass. Multiply each subscript in the empirical formula by this number, and approximate.
Easy Example of Empiricalness	(graciously provided by Howe public schools) Determine the empirical formula of a compound that is composed of 36.5% sodium, 25.4% sulfur, and 38.1% oxygen. 1. Since the amount of each element is given in percentage, you must convert the percentage to a mass. If 100 grams of the sample are assumed, the percentages given are the same as grams. $\text{Moles of Na: } 36.5\text{ g Na} \cdot \frac{1\text{ mole Na}}{23\text{ g Na}} = 1.59\text{ moles}$ $\text{Moles of S: } 25.4\text{ g S} \cdot \frac{1\text{ mole S}}{32\text{ g S}} = .79\text{ moles}$ $\text{Moles of O: } 38.1\text{ g O} \cdot \frac{1\text{ mole O}}{16\text{ g O}} = 2.38\text{ moles}$ 2. Divide each mole number by the smallest mole number. $\text{Na}: \frac{1.59}{.79} \approx 2$ $\text{S}: \frac{.79}{.79} \approx 1$ $\text{O}: \frac{2.38}{.79} \approx 3$ 3. The empirical formula is Na_2SO_3 Now, we calculate the molecular formula. There needs to be a given number of grams of the substance in order to use this. Say that it is 378 g. We calculate the grams of substance from the empirical formula, and we get 126 g. We divide 378 by 126 and multiply it to the subscripts in the equation, and we get a molecular formula of $Na_6S_3O_9$.

VII. Equations and Reactions

Balancing Equations	This is basically an educated guess-and-check method. Things to look out for: make sure the same number of each atom appears on each side, make sure that all the coefficients and subscripts are integers, and make sure that you don't add any more products or reactants to the reaction.

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Types of Reactions	1. Combination: A + B → C. Two or more reactants combine to form a product. An example is $2K + Cl_2 \rightarrow 2KCl$, or Potassium + Chlorine gas → Potassium Chloride. 2. Decomposition: A → B + C. One reactant decomposes to many products. An example is $2HgO \rightarrow 2Hg + O_2$, or Mercury (III) Oxide decomposes to Mercury and Oxygen. 3. Single Replacement: AB + C → AC + B. An example is $2Li + 2H_2O \rightarrow 2LiOH + H_2$, or Lithium added to water forms Lithium Hydroxide and Hydrogen gas. 4. Double Replacement (or metathesis): AB + CD → AD + BC. The reactants switch partners. An example is $2NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2HCl$. 5. Combustion: Hydrocarbon + Oxygen → Carbon Dioxide, water $C_xH_y + O_2 \rightarrow CO_2 + H_2O$. Only the hydrocarbon changes; the rest NEVER does. 6. Dissociation: $A \rightarrow B^+ + C^-$. This is NOT decomposition, as the substance turns into ions instead of compounds. This is when an ionic substance breaks down; i.e. $NaCl \rightarrow Na^+ + Cl^-$.
More Specific Types	1. Neutralization: Double replacement in which acid + base forms salt and water. 2. Precipitation: When a precipitate (a solid) is formed by a reaction of two aqueous reactants.
Electrolyte	A conductor which is nonmetallic in which a current is carried by the movement of ions. It can also mean a substance in which when dissolved in water becomes conductive.
Activity Series	This is a list in order of reactiveness of different elements. It is to determine whether things will react or not in a single replacement reaction. If the single element is MORE reactive than the compound, the reaction will occur, and single replacement will happen. If it is less reactive than the compound, then nothing will happen. For example, Li is more reactive than Cu. Therefore, $Li + CuSO_4$ will react.
Solubility Rules	This is to determine whether the products of a reaction will precipitate or not. The solubility rules are as follows: 1. All compounds containing alkali metal cations and the ammonium ion are soluble. 2. All compounds containing NO_3^-, ClO_4^-, ClO_3^-, and $C_2H_3O_2^-$ anions are soluble. 3. All chlorides, bromides, and iodides are soluble except those containing Ag^+, Pb^{2+}, or Hg_2^{2+}. 4. All sulfates are soluble except those containing Hg_2^{2+}, Pb^{2+}, Sr^{2+}, Ca^{2+}, or Ba^{2+}. 5. All hydroxides are insoluble except compounds of the alkali metals, Ca^{2+}, Sr^{2+}, and Ba^{2+}. 6. All compounds containing PO_4^{3-}, S^{2-}, CO_3^{2-}, and SO_3^{2-} ions are insoluble except those that also contain alkali metals or NH_4^+. Visit http://www.ausetute.com.au/solrules.html for a huge table on these. It's quite pretty too.
Gas Splint Tests	Splints are used to determine if there is a gas produced by a reaction. To test for hydrogen, put a lighted splint into the test tube. If there is hydrogen, it will pop. To test for oxygen, put a non-lighted but glowing splint into the test tube. If there is oxygen, it will reignite.

VIII. Moles

Definition	A **mole** is defined as the amount of a substance containing the same number of molecules, atoms, or formula units as the number of atoms in 12 grams of C-12, which is $6.022 \cdot 10^{23}$.

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Avogadro's number	$6.022 \cdot 10^{23}$. The number of molecules, atoms, or formula units in a mole.
Calculating things	Use factor-label to calculate the answer. For example: how many atoms are in 5000. g of carbon? $5000.g \cdot \frac{1\text{ mol}}{12.011\text{ g}} \cdot \frac{6.022 \times 10^{23}\text{ atoms}}{1\text{ mol}} = 2.507 \times 10^{26}$ atoms. Remember to follow the rules of factor-labeling. Also remember: When there are, say, 5 moles of H_2O, remember that when you split them into atoms, there are 5 moles of every part, meaning that there are 10 moles of *H* and 5 moles of *O*.
Molar Mass	This is defined as the mass of one mole of the substance. This is equivalent to the atomic mass on the periodic table; i.e. hydrogen's atomic mass is 1.00794, so its molar mass is 1.00794 grams per mole.
Molarity	The molarity of a substance immersed in a solution is how many moles of it occur in 1 L of solution. For example, if you have 6-Molar HCl, for every 1 mL of solution, you will get 6/1000 = .006 moles of HCl.
Problems involving Molarity	For example, how many grams of potassium carbonate (K_2CO_3) are needed to make 200 mL of a 2.5 M solution? The molar mass of K_2CO_3 is 138.2 g. A 2.5 M solution means 2.5 moles per liter, so $\frac{2.5\text{ M}}{1\text{ L}} \cdot \frac{1\text{ L}}{1000\text{ mL}} \cdot 200\text{ mL} = 0.5\text{ moles}$. So, we need $0.5\text{ mol} \cdot 138.2\text{ g/mol} = 69.1\text{ g}$ of K_2CO_3. There are other types of problems, but they all follow the rules for molarity. Remember that molarity = number of moles per liter. Derive everything else from that.

IX. Stoichiometry

Definition	Stoichiometry is defined as "the quantitative relationship between constituents in a chemical substance".
Types of Problems	Mass to Mass Mass to Volume Volume to Volume
Conversions between grams and moles	Grams → Moles Divide the number of grams of the substance by the molar mass of the substance. Moles → Grams Multiple the number of moles of the substance by the molar mass of the substance.
Mass to Mass problems	These problems are generally of the form "If A and B react to form C, and you have x amount of A and sufficient B, how much C will you make?" To do this, you must know how to balance equations, work with molar masses, and convert between grams and moles. To solve one of these problems: 1. Balance the equation. 2. Convert the mass of A into moles by using the molar mass. 3. Look at how many moles of A react to form how many moles of C in the balanced equation, and set up a proportion to solve for the amount of moles of C. For example, if 2 moles of A react with 5 moles of C, and you have 3 moles of A, you will make 3 * 5 / 2 = 7.5 moles of C.

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	4. Using the molar mass of C, convert the moles just calculated a mass of C.
Mass to Volume Problems	These problems are generally of the form "If A and B react (or are the products of) C (gas) and D, given the mass of A, find the volume of C." To do this, follow these steps: 1. Balance the equation. 2. Convert the mass of A into moles using the molar mass. 3. Look at how many moles of A react to form how many moles of C in the balanced equation, and set up a proportion to solve for the amount of moles of C. For example, if 2 moles of A react with 5 moles of C, and you have 3 moles of A, you will make 3 * 5 / 2 = 7.5 moles of C. 4. You now have the number of moles of C; multiply this by 22.4 L to get the volume of C at STP.
Volume to Volume Problems	This is basically the same as mass to volume, but instead of converting the mass of A into moles, convert the volume of A into moles by using the formula Volume = 22.4 L times Moles (at STP). So, Moles = Volume / 22.4.

X. Equilibrium

Definition	A state where the reaction rates for a reversible reaction are equal. Or, when in an equation that happens both ways, when there are the same amount of reactions going one way as going the other.
Le Chatelier's Principle	There are three ways to shift the equilibrium: 1. A change in temperature (Δt) 2. A change in the concentration of the reactants or products 3. A change in the pressure ($\Delta pressure$) #3 can occur only where there are different number of moles on the two sides of the equation.
Equilibrium constant expression	$aA + bB \rightleftharpoons cC + dD$ $K_{eq} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$ [x] means the concentration of x with relation to other chemicals in the reaction. If K_{eq} is < 1, then the reactants are favored If K_{eq} is > 1, then the products are favored 'favored' means there will be more of that substance in a given concentration of the solution in which both reactions are going on.
Factors that affect K_{eq}	1. Change in pressure 2. Change in temperature 3. Change in concentration of the products or reactants

XI. Acid-Base Chemistry

General	Water is amphoteric – it can act both like an acid and base $H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$

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	$$K_{eq}=\frac{[H^+][OH^-]}{[H_2O]}$$ But, amount of H_2O too great in comparison to H^+ and OH^- $[H^+] = [OH^-] = 1\times10^{-7}\,M$ Adding an acid will increase the H^+ concentration. A strong acid dissociates and thus increases the concentration of H^+ ions. Ex.: $HCl \longrightarrow H^+ + Cl^-$ Adding a base will increase the OH^- concentration. Ex.: $NaOH \longrightarrow Na^+ + OH^-$
Properties of acids, bases	*Acids* Chemical • $pH < 7$ • Have a high concentration of H^+ ions • Strong acids completely dissociate in water. Physical • Turns blue vegetable dye red • Tastes sour (ex. vinegar, sour milk and lemon juice) • Neutralizes bases • Conducts electricity • When reacted with a reactive metal, it produces hydrogen gas (H_2). Reactive metals include the alkali metals (Group I, Li to Rb), the alkaline earth metals (Group II, Be to Ra), as well as zinc and aluminum. *Bases* Chemical • $pH > 7$ • Have a high concentration of OH^- ions • Strong bases completely dissociate in water. Physical • Bitter taste • Restore the original blue color of litmus after having been reddened by an acid. • Neutralize acids • Conducts electricity • Feel slippery to the touch
Indicators	*Acids* • Turns blue litmus paper red • In universal indicator, it turns reddish (pink, orange, yellow) • Nothing happens when phenolphthalein is put into it • Turns bromthymol blue yellow *Bases* • Turns red litmus paper blue • Turns universal indicator bluish • Turns phenolphthalein pink • Turns bromthymol blue blue

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pH and pOH meaning	pH – measure of H^+ ions in a particular solution pOH – measure of OH^- ions in a particular solution
pH and pOH relationships (calculations/formulas)	$pH + pOH = 14$ $[H^+] + [OH^-] = 1 \times 10^{-14}$ $pH = -\log[H^+]$ $pOH = -\log[OH^-]$ $10^{-pH} = [H^+]$ $[H^+] = \text{antilog}(-pH)$ [*x*] means the concentration of *x* in the solution.
Strong and weak acids and bases	*Strong acids* • dissociate completely in solution • Some examples are: HCl - hydrochloric acid HNO_3 - nitric acid H_2SO_4 - sulfuric acid HBr - hydrobromic acid HI - hydroiodic acid $HClO_4$ - perchloric acid *Strong bases* • Strong acids dissociate completely in water. • Some examples are: LiOH - lithium hydroxide NaOH - sodium hydroxide KOH - potassium hydroxide RbOH - rubidium hydroxide CsOH - cesium hydroxide *$Ca(OH)_2$ - calcium hydroxide *$Sr(OH)_2$ - strontium hydroxide *$Ba(OH)_2$ - barium hydroxide Completely dissociated in solutions of 0.01 M or less *Weak acids* • Weak acids do NOT dissociate 100% in water. (usually anywhere from 1 to 5%) • Weak acids include almost all of the acids except those in the above list of strong acids. • An example of a common weak acid: acetic acid *Weak bases* • Weak bases do NOT dissociate 100% in water. • Weak bases include almost all of the bases except those in the above list of strong bases. • An example of a common weak base: ammonia (NH_3) For an in-depth explanation, go to: http://dbhs.wvusd.k12.ca.us/AcidBase/pH-Strong-Acid-Base.html
Acid rain	Sulfur dioxide (SO_2) and nitrogen oxides $(NO)_x$ undergo reaction with water vapor (H_2O) and other chemicals to yield sulfuric acid (H_2SO_4), nitric acid (H_2SO_4), and other pollutants called nitrates and sulfates. The sulfur dioxide and nitrogen oxides are released as waste from vehicles,

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	factories, etc. It then travels and condenses in the atmosphere to form sulfuric acid, etc, and falls to the ground as precipitation.
Titration problems	Ex.: If 20 cm3 of a 0.3 M solution of NaOH is required to neutralize 30.0 cm3 of a sulfuric acid solution, what is the molarity of the acid solution? *Solution* 1. Write a balanced chemical equation: $2\ NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$ 2. Determine the molarity of the H_2SO_4 $\left(20\ cm^3 \cdot \frac{1\ mL}{1\ cm^3} \cdot \frac{1\ L}{1000\ mL} \cdot \frac{.3\ mol\ NaOH}{1\ L}\right) \cdot \left(\frac{1\ mol\ H_2SO_4}{2\ mol\ NaOH}\right) \cdot \left(\frac{1}{30\ mL}\right) \cdot \left(\frac{1000\ mL}{1\ L}\right)$ $= 0.1 \frac{mol}{L} = .1\ M$
K_w	This is the ionization constant of water or other solutions. It is defined as $[H_3O^+] \times [OH^-]$. It is affected by temperature; when it is very low, water does not ionize very well; when it is high water ionizes very well.

XII. Gases

Kinetic Molecular Theory	1. Kinetic energy depends on the mass and velocity of gas molecules $K_e = \frac{1}{2} mv^2$ *The greater the mass, the greater the energy* *The greater the velocity, the greater the kinetic energy* 2. Gases are tiny particles or molecules with point masses of negligible size 3. Gas molecules are in constant random motion, colliding with each other and with the walls of the container 4. These collisions are elastic – no kinetic energy is gained or lost by the particles in collisions 5. The average kinetic energy of the molecules is the Kelvin (K) temperature of the sample
Diffusion rates of gases	This is how fast gases diffuse, or how fast they flow from a higher concentration to a lower concentration. It is very important in figuring out the movement of gases in soil, air, etc.
Pressure measurements with barometers and manometers	*Barometers* Measure atmospheric pressure *Manometers* Smaller samples of gas Open-ended – gas measure against atmospheric pressure Close-ended – direct measurement of gas sample
Gas Laws	*Boyle's Law:* $P_1V_1 = P_2V_2$ This means that as pressure increases at one rate, volume also increases at that rate. *Gay-Lussac's Law:* $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

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	This means as temperature increases at one rate, pressure also increases at that rate. *Charles's Law:* $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ This means as volume increases at one rate, temperature also increases at that rate. *Combined Gas Law:* $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ This means that the rate of change of temperature is equal to the rates of change of pressure and volume. This is better described in the Ideal Gas Law. P = pressure V = volume T = temperature
Ideal Gas Law	$PV = nRT$ P = pressure V = volume T = temperature n = number of moles R = universal gas constant $\left(\frac{8.31\ \text{kPa}\cdot\text{L}}{1\ \text{mol}\cdot\text{K}} \text{ or } \frac{.08205\ \text{atm}\cdot\text{L}}{1\ \text{mol}\cdot\text{K}}\right)$ You don't need to know the universal gas constant, but this law can be summarized as follows: Pressure and Volume are directly proportional to Temperature. So if P increases, the T increases *or* V decreases to keep the equation balanced.
Molar Volume of a Gas	At STP, it is 1 mole of gas = 22.4 L of gas
Definition of STP	"Standard Temperature and Pressure". T (Temperature) = 0^{o}C P (Pressure) = 1 atm (atmosphere) = 760 mmHg = 101.3 kPa = 14.7 psi = 760 Torr
Avogadro's Hypothesis	Equal volumes of gases at the same temperature and pressure contain equal numbers of particles. Ex: $2\ C_2H_6 + 7\ O_2 \longrightarrow 4\ CO_2 + 6\ H_2O$ $18.2\ L\ C_2H_6 \times \frac{4\ L\ CO_2}{2\ L\ C_2H_6} = 36.4\ L\ CO_2$
Dalton's Law of Partial Pressures	$P_{total} = P_1 + P_2 + P_3...$ Pressure of the whole gas mixture in a sealed container is equal to the sum of the pressures of each gas in the mixture. i.e. a mixture of oxygen, neon will have pressure of (pressure of oxygen) + (pressure of neon).

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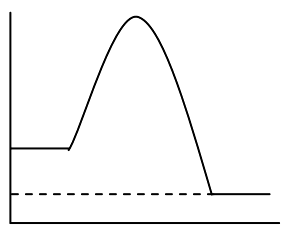
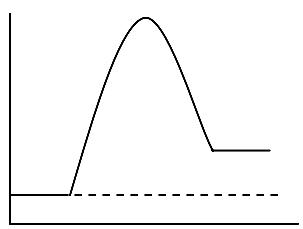
XIII. Thermodynamics	
Endothermic reactions	Endothermic reactions absorb energy ςH is positive energy is written as a reactant
Exothermic reactions	Exothermic reactions release energy ςH is negative energy is written as a product
Specific heat / specific heat capacity	Specific heat – amount of energy a substance needs to gain in order to change its temperature Specific heat capacity – amount of energy required to change the temperature by $1^{o}C$ for 1g of a substance Units are in $\frac{\text{Joules}}{g \cdot {}^{O}C}$ or $\frac{\text{cal}}{g \cdot {}^{O}C}$ specific heat capacity for H_2O = $\frac{4.186 \text{ Joules}}{g \cdot {}^{O}C}$
q = mCςt	q = mCςt q – ςH = enthalpy C – specific heat (for that particular substance) ςt – change in temperature
Potential energy diagrams	*Exothermic reactions:*  Left: reactants Right: products Difference between right and left: energy released/absorbed *Endothermic reactions:* 
Hess's Law	Add reactions. Add products. Add enthalpy. (manipulate the given equations to form the equation whose enthalpy you are trying to find)
Other	ςH°f = ςH°P - ςH°R

This is the end of the study guide. If you find any errors or have any questions or comments about this study guide, feel free to email me at fenguin@gmail.com. Thanks a lot for reading, and good luck on finals!

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