Thank you being a reader of this blog. Any feedback you wish to provide would be appreciated. You can email comments or talk to me in class. You can also look for a survey soon coming by email that you can fill out. This survey is completely anonymous. The goal is to keep this blog and podcast going throughout the term in a way that is most useful to everyone. I plan to use this forum to introduce chapters and offer post-chapter reviews of chapters. Some answers to FAQs can also be a focus of posts. Any suggestions for topics to discuss are encouraged.

Today's blog post and podcast are on the topic of useful topics from high school chemistry for review prior to beginning Chem 1210. Many of the topics of Chemistry 1210 will likely be things you've already learned about in high school. AP chemistry is often very similar in content to our class and even regular high school chemistry covers most of the main topics often though with a little less depth. One particular way Chem 1210 might be more advanced than these high school courses is that we have a strong focus on the ideas and concepts that you must understand for success in later courses, like Chem 1220, Chem 2310/2510/2520/2540/2550 (organic chemistry courses), and upper level biochemistry, genetics, and materials science courses.

As a means to review and/or introduce some of the main topics for the course, this blog post will review some of the key terms and ideas from the first 11 chapters of our textbook.

Chapter 1:
Chemistry: Properties and reactions of matter.

Chemical reactions: transformation of a substance or substances to new substances. We will discuss what a chemical substance means and definitions of elements, compounds, and molecules in chapter 1.

Significant figures: One of the more challenging concepts of how to apply, but this system helps know how to round results of mathematical calculations involving measurements. Measurements always have some fundamental limitation on how well they are known. Like a scale in your bathroom might weight out to the nearest 1 pound or maybe to the nearest 0.1 pound. Most objects could be described by a much more precise mass than this, though this bathroom scale would be limited to only the placeholders shown on its display. Often you find tools this have some uncertainty in that last placeholder – you might see some fluctuation and you would try your best to take the mass as the average of these values when possible. When you do a calculation on a number like 123, you ordinarily would think this means 123.00000000 but if 123 represents the number of pounds read from a scale, it would really mean 123 ± 1. The true mass of the object is very likely between 122.5 and 123.5 and we cannot know its mass any better than this range. This fundamental limitation will limit the number of placeholders that are significant in calculations of this value, like addition/subtraction, multiplication/division, natural logs, etc. We will go over these rules in the first week of lecture.

Precision vs accuracy of a measurement: Precision is the repeatability of a measurement and accuracy is how close a measurement is to a true value. Accuracy can only be tested by comparison to a known value, like using an exactly 1.00000±0.000001 kg object to test the mass of a balance. The topic of significant figures has more to do with precision of a measurement, in which a more precise measurement is one that contains more significant digits. A mass recorded on a balance of 125.5 pounds is more precise than another balance which reads its mass to be 126 pounds. Both measurements are essentially just as accurate as each other though. We either test a tool in the lab or we will just assume a given tool is accurate its given number of placeholders.

Rules of significant figures, which we will discuss thoroughly in lecture, are generalized rules for statistical analysis when dealing with measured quantities that you might learn more about in an introductory stats class. Significant figures are a key item of discussion in chapter 1 so we will explore this topic in lecture, so this is not a topic I expect most of you to have much knowledge about from high school chemisty.

One final topic from chapter 1 is dimensional analysis. You likely recall that this method is often used to convert units for problems like figuring out how many seconds are in 1 year. We can also use dimensional analysis as a general means of solving problems any time equalities are known. One equality that can easily be used in dimensional analysis is density, like the density of water is 1.00 g/mL. If you have 55.0 mL of water, it would of course have a mass of 55.0 g because the density says you have 1.00 g per every mL. We hardly need any method to solve that problem. But what if we're talking about ethanol, which has a density of 0.789 g/mL, what is the mass of 55 mL of ethanol? Well, now we might right this out more carefully:

# g = 55 mL x 0.789 g /mL = 43.4 g

All we need to know to set this problem up is that the mL must cancel out, so we must therefore multiply by the density and then this calculation results in our desired unit of grams being in the numerator. We could have instead used the formula for density, which is d = m/V, rearranged for m, and plugged and chugged. This problem solving method is totally correct too, but we did have to remember an equation to use that method and we didn't have to remember much of anything to use dimensional analysis. We will often use dimensional analysis in this course to solve a variety of problems in ways that ought to make the solutions simpler than other methods.

I didn't mention this detail in the podcast, but there is really never anywhere on exams or recitations where you have to show work (recitations are just attendance points and exams are all multiple choice). So how you solve your homework problems is mostly irrelevant, provided you ultimately get the answer right! I always recommend though that you write out units in your work though as this process will help eliminate mistakes.

Topics in chapter 2 include modern atomic theory, atomic and molecular weights, and some examples of common compounds and rules for naming classes of compounds. The first two sections discuss theories about the discovery of the subatomic particles like the electron (negative particle), proton (positive particle), neutron (neutral particle), and how the theory of the nuclear atom was discovered. We will briefly discuss the periodic table and it's structure based on increasing size and grouping of elements by similar properties.

Each of the elements differ by the number of protons within the nucleus of the atoms comprising the element. Hydrogen atoms have 1 proton in their nucleus, carbon atoms have 6. The number of protons hence defines the identity of the atom so it's called the atomic number. The neutron is a neutral particle and atoms of some elements can differ in how many neutrons they contain. Atoms of the same element with different number of neutrons are isotopes. Carbon for example most commonly exists as carbon-12 and carbon-13. The number after the element is the mass number, the sum of the protons and neutrons (so carbon-12 has 6 protons and 6 neutrons, about 99% natural abundance of carbon, and carbon-13 has 6 protons and 7 neutrons, which is about 1% natural abundance of carbon). These isotopes often behave very similarly and they are only different by the small difference in mass. Electrons are not counted in the mass number primarily because their mass is very low, around 1/1000th the mass of a proton or neutron. The electrons being so light are one reason why they do not coexist among the protons and neutrons but instead spin around the positively charged nucleus. The proton and electron have the same magnitude of charge, just opposite in charge. So a neutral carbon atom contains 6 electrons. A fluorine atom (9 protons) with 10 electrons has a charge of -1 and hence this is an ion. Na^+ is the ion of sodium with 11 protons and 10 electrons. We will see often atoms near the ends of the periodic table can lose electrons (on the left side) or gain electrons (on the right side) so that they arrive at a stable noble gas count of electrons. We will see later in chapter 6 how electron configurations and shells play a role.

Metallic elements only lose electrons to form cations and they will never gain electrons and form anions. Metallic elements as we will see in chapter 7 have low ionization energies. Sodium's group (the alkali family) will lose one electron to form +1 cations, calcium's group (alkaline family) will lose two electrons to form +2 cations. The transition metals can often lose a variable number of electrons that is not very predictable, though they would never lose any electrons past the nearest noble gas count (for example, titanium can form the Ti^{4+} cation as its highest charged cation since it would have 18 electrons like argon). In order for a metal (or any atom) to lose electrons, another atom must be able to accept those electrons. If Na(s) and Cl_2 (g) were allowed to react together, two sodium atoms would lose one electron each and the Cl atoms in a Cl_2 molecule would gain one electron each and the compound NaCl would result. This compound is an ionic compound, Na^+ and Cl^- ions would be present and their electrostatic attraction keeps the salt crystal held together.

We will see common compounds and naming conventions in the end of chapter 2. We will also see some basics of formulas and charges of ions in common ionic compounds.

Chapter 3 deals with chemical reactions, such as writing, balancing, and solving problems related to basic chemical reactions. A chemical reaction shows the reactants on the left side and the new substances that result in the products. A chemical reaction is not really an equality, rather it is much better thought of as a 'before and after'. A chemical reaction never changes the identities of the elements, rather only new compounds are formed from rearranging the atoms into new substances in the products. As an example, CH_4 (methane) reacts with O_2 (oxygen gas) to form carbon dioxide (CO_2) and water (H_2O). All of the carbon in methane will end as all of the carbon dioxide (CO_2). All of the hydrogen in methane ends up in the water molecules that form. If 1 molecule of methane are reacted, then 1 molecule of CO_2 and two molecules of H_2O must form to balance the C atoms and H atoms. Two molecules of O_2 must therefore be present in the reactants. Reactions must always have an equal number of atoms in the reactant and product sides of the reaction:

$$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$$

Subscripts are never changed to balance a reaction. Subscripts relate the number of atoms of an element in a molecule or unit of the compound. An oxygen molecule contains two oxygen atoms. “2 O_2” in the above reaction implies two oxygen molecules. O4 would be a totally different molecule, which does not even exist. A few other molecule examples to get a sense of the complexity other molecules have, methanol is CH_3OH, you could write this as CH_4O but CH_3OH is implying something about the molecular structure of the molecule that we will later see more about in chapters 8 and 9. Methanol is a molecular compound and a sample of methanol would contain many molecules of methanol. If you were to boil methanol, individual molecules of methanol would escape from the liquid surface into gaseous vapors (just like water molecules evolve from boiling water). Sulfate ion is SO_4^{2-}. Potassium sulfate is K_2SO_4 (and this is an example of an ionic compound, with an overall neutral compound comprised of ions).

Stoichiometry is another key topic chapter 3, which is the bookkeeping system for reactions. So say if 3.0 grams of methane reacts with excess oxygen, stoichiometry will help us figure out how much CO_2 and H_2O would form. The mole is a common unit we will use. A mole of an object refers to 6.02214179x10^{23} of that object so a mole of methanol contains 6.02x10^{23} methanol molecules (rounding to three sig figs). The periodic table lists the average atomic weights of all of the elements, often rounded to 4-5 sig figs but you can look more sig figs up in the CRC if ever needed. The average atomic weight accounts for the abundance of each isotope of an element and its atomic weight and it sums up all of these quantities for the average atom of each element. For carbon, carbon-12 is defined to have a mass of exactly 12 atomic mass units (amu), so think of its mass as 12.000000000000····· amu. More on this system in a moment, but the mass of carbon-13 is about 13 amu (we will see a precise example in class). If you recall carbon-12 is about 99% natural abundance and carbon-13 is about 0.01, so take 0.99 times 12 plus 0.01 x 13 = 12.01. So the average atomic weight of carbon is just a little over the mass of the isotope that ~99% of the atoms have a mass of, that makes sense I hope! To get a sense of a more precise atomic weight, we’d have to plug in the exact mass and abundance of the carbon-13 isotope. Back to the amu acale, so carbon-12’s mass was defined to be exactly 12 amu. If you take as exact mass of 12 grams of carbon-12 as you can find a balance to measure (about 9 sig figs is possible) and start counting atoms, you’d come up with about 6.02214179x10^{23} atoms (Avogadro figured this out). What this means is that each atom weights 12.0000000 g / 6.02214179x1023 atoms = 1.992···x10^{-23} g. This then also means that 12 amu = 1.992x10^{-23} g, or that 1 amu = 1.660····x10^{-24} g, or that 1 g = 6.022···x10^{23} amu. OK, if you can follow all of that, that’s great! We will talk about that derivation in class, but there is a simple conclusion we can either draw from this or just take for granted, that is that the periodic table tabulated the masses of each atom, either in the unit of amu per average atom or in grams per mole of the element. So oxygen on the periodic table says 15.9994. This means either an average oxygen atom has a mass of 15.9994 amu or that one mole of oxygen has a mass of 15.9994 g. We could call the molar mass of oxygen to 15.9994 g/mol. We can then use the periodic table to determine molar masses of compounds just by adding up all of the average molar masses of its elements, so CH_4 = 12.01 + 1.008*4 = 16.04 g/mol.

Going back to the $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$ reaction, the reactants and products can be related in terms of molecules or moles: 1 molecule of CH_4 reacts with 2 molecules of O_2 to form 1 molecule of CO_2 and 2 molecules of H_2O OR 1 mole of CH_4 reacts with 2 mole of O_2 to form 1 mole of CO_2 and 2 mole of H_2O. If 3.00 grams of CH_4 reacts with excess O_2, and we wanted to know how much CO_2 forms, we could convert grams of CH_4 to moles, moles of CH_4 can then be converted to moles of CO_2, and then the molar mass of CO_2 (44.01 g/mol) can be used to convert moles of CO_2 to grams of CO_2; 8.23 grams of CO_2 would form.

Stoichiometry also gets into which reactant limits the amount of product formation and which might be present in excess, or what the percent yield is of a reaction. Chapter 4 gets into reactions taking place in water, such as acid-base reactions and metathesis reactions (which are reactions where you take two water-soluble ionic compounds, like NaCl (aq) and $AgNO_3$ (aq) when mixed lead to AgCl forming, which is a white solid, which would appear to form as snowflakes in the solution and ultimately settle on the bottom of the reaction flask. AgCl forms as a solid because the compound is insoluble in water. Solubility trends of ions will be discussed so we can product when metathesis reactions will occur. Chapter 4 also discusses oxidation-reduction reactions, which are reactions where electrons are gained (i.e. reduction) and lost by atoms (i.e. oxidation) during the reaction. Chapter 4 wraps up with solution stoichiometry problems, such as titrations.

Chapter 5 discusses heat capacity, like how much heat must water absorb to raise its temperature, and enthalpy changes of chemical reactions. A later chapter in Chem 1220 would get into other thermodynamics topics.

Chapter 6 deals with properties of electrons in atoms, starting with the hydrogen atom in greater detail and then electron configurations for larger atoms towards the end of the chapter (e.g. you might recall the ground state electron configuration of carbon is $1s^2\ 2s^2\ 2p^2$).

Chapter 7 deals with nuclear and periodic trends. One such periodic trend is ionization energy, which mostly increases from left to right across a row of the periodic table but decreases from top to bottom down a group. We will discuss why, as well we will discuss other trends like atomic size, ionic size, and electron affinity trends. These trends can help us understand simple bonding trends in ionic compounds, like sodium should has a relatively low ionization energy, so not much energy is required for the atom to lose an electron. Chlorine on the other hand has a much higher ionization energy. So if you are considering the ionic structure to be either Na^+Cl^- vs. Na^-Cl^+, the latter would take more energy to form Cl^+ than it would take for sodium to lose an electron, so $Na^{-}Cl^{+}$ is unlikely to form with those charges. Additionally, the electron affinity of Cl is more negative, as electron affinity decreases (becomes more negative and hence more favorable for an atom to absorb an electron) from left to right. So Cl is energetically more favored to accept an electron than Na. So we can conclude that NaCl forms the structure we have thought all along, $Na^{+}Cl^{-}$.

Chapter 8 starts with a description of ionic bond strengths as a relationship to the lattice energy of the ions. As ions are more highly charged in magnitude, the lattice energy increases hence the ionic bond strength is stronger, and then an associated trend like melting point should increase. For example, we can compare NaCl (mp ~ 800 ˚C) to MgO (mp 2850). MgO has higher charged ions of $Mg^{2+}O^{2-}$, so the attraction of these ions is much greater. The compound has a lattice energy about four times that of NaCl and likewise its melting point is much greater.

The second half of chapter 8 is centered around Lewis structures of molecular compounds. Molecular compounds result between non-metallic elements like C, H, N, O, Cl, etc. These elements by nature have higher ionization energies and many have less negative electron affinities. So these compounds are not able to gain and lose electrons like metals and non-metals are able to. Instead, these elements must share electrons with each other in order to gain stability, through bonds called covalent bonds. Covalent bonds are not as strong as ionic bonds. A Lewis structure is a depiction of all of the bonds and bond types (such as single, double, triple, etc) in a molecule. Hydrogen molecule, H_2, has a single bond between its two H atoms, H–H. Each H atom has two electrons near it, allowing each to have an electron count like He – it's nearest Noble gas element. Two oxygen atoms, each with 6 valence electrons, can adopt a configuration of 8 valence electrons by forming the oxygen molecule with a double bond: O=O. CH_3OH for example has three single bonds between C and H, a single bond between a C and O atom, and a single bond between O and H. Through some simple rules and ideas, we can get into some bigger molecules toward the end of chapter 8.

Electronegativity and bond polarity is discussed in chapter 8 and we discuss molecular polarity in chapter 9. We also talk about three dimensional molecular shape in chapter 9. Chapter 9 also discusses molecular shape. Organic chemistry builds off many of the ideas from chapters 8 and 9.

Chapter 10 is on the topic of gases and begins with the ideal gas law. In this law, gases are described as non-interacting particles. Not much chemistry happens between non-interacting particles! However, all gases will eventually condense, hence their particles must at some low temperature begin to interact. These interactions can be described by the van der Waals equation toward the end of chapter 10 and the subsequent liquid state is discussed in chapter 11.

Chapter 11 discusses many important factors of liquids, such as intermolecular forces of dispersion forces (force present in liquids of nonpolar substances), dipole-dipole forces (force present in liquids of polar substances), hydrogen bonding (force present in liquids of polar substances where H is directly bonded to an N, O, or F atom). These intermolecular forces are weaker than covalent bonds but they are still very important for understanding properties of liquids. Many properties and trends of a liquid can be inferred from its types of intermolecular forces, and these forces can be implied from knowledge of the shape and structure of a molecule. Chapter 12 completes a discussion of the phase of substances by discussing solids.

This is much longer than I thought it would be! I hope this gives a sense of what this course is about and that you probably have heard many of these words and ideas before. We will expand upon these ideas and focus on problem solving in our lectures.

There is another handout you can read on some of these topics. It’s posted below on the blog – check it out if you want to read more.

This is the last blog entry of topics I had planned before our class starts up. If you have questions, keep them coming as I could do more posts on FAQs if there are still questions. I also plan to keep the blog going throughout the term though I’m exactly sure how I plan to do this. My plan at the moment is I could provide pre and/or post chapter summaries to help introduce topics and to help review material. If you have ideas for future blog topics, let me know!