

## **Key concepts of chemistry: What you should know before Chemistry 1210!**

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**Introduction to college chemistry:** The transition from high school to college can be challenging. College courses are often much less structured than high school courses. Exam scores are a higher percentage component of your grade than in high school, as well exams tend to cover more material than in high school. Most college courses will have cumulative final exams that can be worth 20-40% of your final grade depending on the course. Material is covered at a faster pace in college with little to no class time spent on review. It is truly to your advantage to use a learning style from the first day of class that is conducive to long-term learning. Don't plan to cram your studying into the night before a test and instead focus on productive habits that will make you successful in this class and beyond!

I have observed good and bad study habits over the past six years as an instructor of introductory chemistry courses at The Ohio State University. Traits in my successful students are attending lecture and recitation, reading the book before and/or after class, figuring out how to solve the homework problems, and doing plenty of practice exam problems. Interestingly, many of my most successful students tell me they don't necessarily always do those things, but more importantly that they can tell when they understand a concept or not. Assessing your strengths and weaknesses is perhaps the most important skill to develop to become a successful student in the physical sciences.

Whether or not this is your first semester at college or not you will probably agree: It's easy to skip classes since attendance is not required or even recorded in most classes; It's easy to put off reading the book before or after class because most classes do not have weekly exams or high-value quizzes; It's easy to passively do homework and practice exam problems (like looking up a solution to a problem online or in a solution manual or you have a friend/TA/instructor work a problem for you instead of authentically solving it using your own knowledge and problem-solving skills). The simple truth is the knowledge and skills you develop by attending class, reading the book, and working homework and practice exam problems directly translates into the skills you will have to use during exams. So come to class, read the book, and work homework as much as you need to become good at solving exam problems! Through the 15+ practice and sample exercises within each chapter, the 100+ homework problems following each chapter, and the 50-100 practice exam problems I offer through our course website, you should have plenty of assorted exercises to help prepare you for our exams.

Yet another challenge is that most college courses expect you have come with essential skills from prior high school classes. For Chemistry 1210, the key prerequisite courses are high school chemistry and trigonometry. If you tend to struggle with math-word problems, you may seek a math problem-solving book, like an ACT or SAT study guide. To give a better sense of the chemistry topics we hope you remember from high school, I have prepared this review of some of the basic principles of chemistry which hopefully are a review, though you will see we will discuss, use, and build from these concepts throughout Chemistry 1210, as well in Chemistry 1220, 2210, 2510, 2520, 2540, and 2550.

**Atoms, elements, and ions:** One of the most fundamental concepts in chemistry is that the atom is the fundamental building block of matter. The atom, once thought to be the least divisible form of matter, is comprised of three key subatomic particles. According to modern atomic theory, an atom contains protons and neutrons within a compact nucleus and electrons spin around the nucleus. **Protons** are positively charged particles, **electrons** are negatively charged particles, and **neutrons** have no charge. The charge of a proton is equal to the magnitude of the negative charge of an electron, hence a neutral charged atom must contain an equal number of protons and electrons. We can calculate the charge of an atom or ion as the number of electrons subtracted from the number of protons:

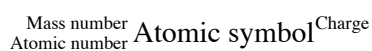
$$\text{Charge of atom} = \# \text{protons} - \# \text{electrons} \quad (\text{Eq. 1})$$

The mass of a proton and neutron are similar though not exactly equal, yet both are much more massive than the electron. As a result, the electron has a very high velocity and it is found to spin around the much more massive nucleus. The orbital nature of electrons will be discussed in more detail in chapter 6. Nuclear and periodic trends will be discussed in chapter 7.

The **periodic table** of elements shows over 100 elements that have been discovered. What differentiates elements from each other is the number of protons contained within the nucleus of an atom of that element. The **atomic number** defines the number of protons in the nucleus of an atom, hence identifying the element. An atom of hydrogen, the lightest element, contains just 1 proton. An atom of the next heaviest element helium contains 2 protons. The periodic table is arranged by increasing number of protons, as well the periodic table is arranged into groups of atoms in vertical columns that share a common set of properties.

The atoms of elements often contain various counts of neutrons within their nuclei. For example, hydrogen atoms can be found in nature with either zero neutrons (~99% of the atoms) or with one neutron (<1%) within its nucleus. Likewise, carbon atoms are found to contain either 6 neutrons (~99% of atoms), 7 neutrons (~1% of atoms), or 8 neutrons (<1 in 10 billion carbon atoms). Atoms that share the same number of protons yet differ in number of neutrons are termed **isotopes**. Isotopes of an element have very similar properties since they differ only in the number of neutral-charged particles. A carbon with 6 neutrons behaves very much the same as a carbon atom with 7 neutrons. Atoms smaller than lead (Pb) generally are stable, so the proton and neutron counts for the compounds we will handle in the lab are effectively constant, at least on the time scale of our lifetimes!

The protons and neutrons represent a majority of the mass of an atom. The sum of these particles is called the **mass number**. Carbon-12 and carbon-13 have mass number of 12 and 13, and hence 6 and 7 neutrons, respectively. The **atomic symbol** is also a way to show an atoms mass number:

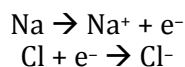


A carbon atom with 6 protons, 6 electrons, and 7 neutrons can be represented as  $^{13}_6\text{C}$  or  $^{13}\text{C}$ , the latter is more often used since specifying both the atomic number and elemental symbol would be redundant. Since the atom in this example has an

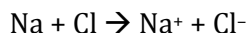
equal number of protons and electrons, the charge of the atom is zero. Only a non-zero charge needs to be specified in its symbol. As an example, a chlorine atom with 17 protons, 18 electrons, and 18 neutrons is represented by  $^{35}\text{Cl}^-$ .

An **ion** is an atom bearing a charge due to an imbalance in its number of protons and electrons. Since the very light electrons spin about the nucleus, electrons can be gained or lost by atoms much more easily than protons. Using equation 1, a negatively charged ion called an **anion** refers to a negatively charged ion with more electrons than protons and positively charged ion called a **cation** refers to a positively charged atom with fewer electrons than protons.

Sodium and chlorine are examples of elements that commonly form cations and anions, respectively. An electron can be lost from sodium yielding  $\text{Na}^+$  and an electron can be added to Cl to yield  $\text{Cl}^-$ :



Sodium above is **oxidized** since **oxidation** represents a reaction where an atom has lost electron(s). Chlorine is **reduced** since **reduction** is represents a reaction where an atom gains electrons. In nature, oxidation and reduction reactions must occur together simultaneously because matter can neither be created nor destroyed. We will more thoroughly discuss the oxidation-reduction chemical reaction in chapter 4.



Some examples of **periodic properties** of the elements are the left side elements on the table are metals, the right side elements are nonmetals, and the elements in between are metalloids (see Fig. 2.15 in your book). Metals commonly form cations and they do not form stable anions. Non-metal atoms mostly form stable anions, though hydrogen is readily oxidized to the  $\text{H}^+$  cation.

The **Noble gases** (He, Ne, Ar,...) are a group of stable elements. These elements are stable because they have a completely filled shell of electrons. We will discuss more about shells in chapter 6. The **halogen** group (F, Cl, Br, etc) form anions of  $-1$  charge, whereby these anions arrive at the same number of electrons as their nearest respective Noble gas element. The **chalcogen** group (O, S, Se,...) form stable anions of  $-2$  charge and nitrogen's group forms stable ions of  $-3$  charge.

The **alkali metal** group (Li, Na, K, etc) are all very reactive with water and each forms ions  $+1$  charge. The **alkaline metal** group (Mg, Ca, Sr, etc) are not as reactive with water as the alkali metals and they form  $+2$  charged cations. The **transition metals** (Sc-Zn groups) are mostly not reactive with water and can be oxidized to cations of various charges. Many transition metals, like Fe, Cr, and Co can form stable cations of more than one charge. Iron, for example, forms stable ions of both  $+2$  and  $+3$  charge.

The **metalloid** elements (B, Si, Ge, As, Sb, Te, At) have properties of metals and nonmetals. The properties of some of these elements are discussed in chapter 12.

The periodic table lists the **average atomic weight** (AW) of elements, which is defined as the *weighted average* of all natural-occurring isotopes of an element. Isotopes of an element are of differing mass due to their different number of

neutrons. The average AW can be determined by adding up the fraction of atoms that exist as a particular isotope multiplied by the atomic weight of each isotope:

$$\text{Average AW} = \text{fraction isotope a} \cdot \text{AW a} + \text{fraction isotope b} \cdot \text{AW b} + \dots$$

Most periodic tables, like the one in your textbook, list the average atomic weight of each element to several decimal places, though often these masses are more precisely known than this and could be looked up if ever needed in an index like the CRC Handbook of Chemistry and Physics.

**Tip:** You can access the CRC via the OSU library page. A convenient way to access the library page is to login to your Carmen account, enter our course page, and click the library link. This will take you to a page with several links useful for our class. The CRC handbook is an excellent source for melting points, boiling points, solubility information, and much more. Sometimes you need to check it's symbol index to make sense of the information provided.

The **atomic mass unit** scale is devised because of the extremely small size and low mass of atoms. In a 12.0000000 g sample of pure carbon-12, there are precisely  $6.02214179 \times 10^{23}$  atoms. The mass of a single carbon atom is therefore  $1.99269346 \times 10^{-23}$  g. Obviously, this precise, small number would be peculiar to list on a small periodic table, and equally odd to use in problem solving. The atomic mass unit scale has a simple relationship to mass, 12 a.m.u. is *defined* to be equal to *exactly*  $1.99269346 \times 10^{-23}$  g (or the mass of a single carbon-12 atom), which then simplifies into 1 gram being equal to  $6.02214179 \times 10^{23}$  a.m.u.

The number  $6.02214179 \times 10^{23}$  is defined as **Avogadro's number** ( $N_A$ ). Avogadro's number also defines the quantity called the **mole**, where there is an Avogadro's number of particles represents a mole of the particle. So, there is a mole of carbon atoms (i.e. an Avogadro's number of atoms) in the exact 12-gram sample of pure carbon-12 *AND* there is an Avogadro's number of a.m.u. in exactly 1 gram. Thus, a single carbon-12 atom has a mass of *exactly* 12 a.m.u. by the definition of this scale *AND* the molar mass of carbon-12 is 12.0000000 g/mole.

The listed average AW of carbon on the periodic table (12.0107) reflects its average mass is slightly heavier because 1% of its atoms in natural samples are carbon-13 and a very tiny fraction are carbon-14 (less than 1 out of 10 billion atoms). Carbon-13 and carbon-14 have atomic weights of 13.0034 and 14.0032 a.m.u., respectively. The average AW listed for carbon corresponds to its molar mass (i.e. 12.0107 g/mole) *OR* its atomic weight (i.e. 12.0107 a.m.u. per atom).

Average atomic weights and molar masses are most often used to calculate the atomic weight and molar mass of a molecule. As an example, methane ( $\text{CH}_4$ ) has an atomic weight of  $12.0107 + 1.0079 \times 4 = 16.0423$  a.m.u. per molecule or 16.0423 g/mol. We will see some examples of how average atomic weights are calculated and more about their use in chapter 2.

**Bonding in molecules and compounds:** Compounds form when atoms of different elements create bonds with each other in a way that makes their atoms more stable. Five examples of types of bonds that are found in nature are Ionic, Covalent, Metallic, Hydrogen bonds, and van der Waals forces. Solid sodium and gaseous  $\text{Cl}_2$  are more stable in their ionic forms of  $\text{Na}^+$  and  $\text{Cl}^-$ , where the ions once formed have a strong stabilizing electrostatic attraction called the **ionic bond**. Table salt,  $\text{NaCl}$  (s) is called an **ionic compound** since it is comprised of  $\text{Na}^+$

and  $\text{Cl}^-$  ions. Water or  $\text{H}_2\text{O}$  (l) is more stable than hydrogen and oxygen, even in the element's most stable states of gaseous  $\text{H}_2$  and  $\text{O}_2$ . Unlike Na, the energy of to ionize hydrogen to form  $\text{H}^+$  is very high in energy, so water therefore does not exist as an ionic compound like  $\text{NaCl}$  (s). Water instead use **covalent bonds** and it is termed a **molecular compound** since it is comprised of  $\text{H}_2\text{O}$  molecules. The individual water molecules in a liquid sample of water are attracted to each other through hydrogen bonds and van der Waals forces. The bonding in solids or liquids of metallic elements is best represented by the metallic bond, which is discussed in chapters 12 and 23.

Ionic compounds tend to have a repeating three-dimensional pattern of cations and anions. Some of the particular structures will be discussed in Chapter 12. Ionic compounds are best represented with an **empirical formula**, which is the simplest ratio of the ions comprising the compound. Molecular compounds tend to have a particular arrangement of the atoms. Molecular compounds are best represented by their **molecular formula**, which is the specific ratio of elements in the compound. **Metallic solids**, such as Fe (s) exists in the solid state as three-dimensional arrays of atoms whereby all atoms share their valence electrons (the electrons past the most stable filled electron shell, or nearest Noble gas element).

Even though  $\text{H}_2\text{O}$ ,  $\text{NaCl}$ , and Fe can all exist as solids, their properties vastly differ from each other based on their differences in bonding type. The nature of why different types of elements form different types of bonds is most related to their differences in ionization and nuclear charge trends. Ionic bonds, for example, arise within atoms that can most easily form cations and anions. We will discuss in chapter 7 that elements are most easily oxidized to cations the further left on the periodic table (such as sodium) and the elements most easily reduced are further to the right on the periodic table (such as chlorine), though a simple explanation is these elements ionize this way is because they are the fewest electrons away from a stable Noble gas number of electrons.

**Covalent bonds** arise among nonmetal elements, such as H, C, O, N, etc. These atoms require greater energy than metals to become ionized. Instead, these atoms find the greatest stability when they share their outer, valence electrons with each other through covalent bonding. A covalent bond describes a bond where two atoms share electrons. The simplest covalent bond arises between two hydrogen atoms. Two hydrogen atoms, each containing a single valence electron, form a stable covalent bond by sharing their two electrons in a covalent bond. The below Lewis structure is a representation of the bond as a single line representing two shared electrons in a single bond:



**Covalent bonds** arise when they allow the atoms participating in the bonds to adopt a more stable electron count than they would have as separated atoms. Atoms that participate in covalent bonding are called molecular compounds or molecules. Covalent bonding and structures of molecular compounds will be discussed in more detail in chapters 8 and 9.

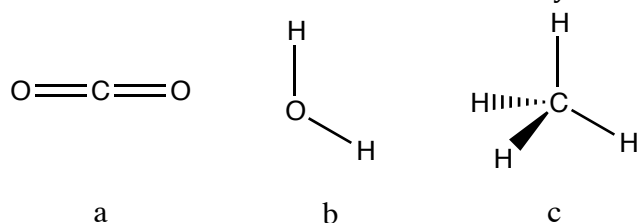
**Metallic bonding** arises in pure samples of metallic elements. Metallic elements often contain several valence electrons and highly charged nuclei. These electrons are best described as being able to flow through the structure of the metal. If you imagine a copper wire, an applied charge (i.e. electrons) causes the electrons in the wire to flow away from this charge due to electron-electron repulsions.

Sometimes the bonding in metals is described as a “sea of electrons”, since there are so many electrons that can effectively freely roam among the nuclei of the atoms, which helps justify why most metals are great electrical conductors.

To summarize thus far, ionic bonds arise when electrons are permanently exchanged between elements leading to oppositely charged ions that are strongly attracted to each other, covalent bonds arise when electrons are discretely shared between atoms, and metallic bonds arise when electrons can freely flow among metallic nuclei. Generally metallic and ionic bonds are much stronger than covalent bonds, which leads ionic and metallic compounds to have higher melting and boiling points than molecular compounds.

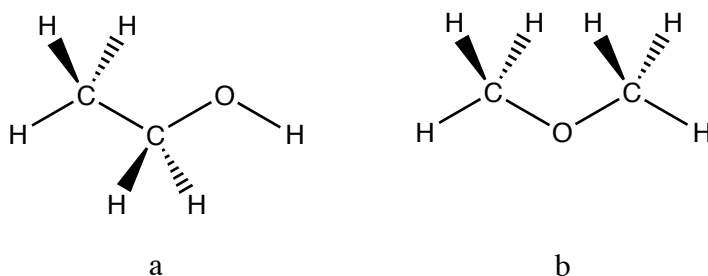
**Hydrogen bonds** and **van der Waals forces** are the attractive forces that exist between molecules. These types of forces are not nearly as strong as covalent bonds, yet these are the forces that allow molecular compounds to be attracted to each other in their liquid and solid physical states. Van der Waals forces are generally weaker than Hydrogen bonds. Since these forces of attraction are weaker than covalent bonds, they are often termed intermolecular forces and/or non-covalent bonds. These forces will be discussed in greater detail in chapter 11.

**Molecular structure:** All molecules have a 3-dimensional shape. A few examples are  $\text{CO}_2$  is a linear molecule, water is a bent molecule,  $\text{CH}_4$  is a tetrahedral-shaped molecule (see Fig. 1). We will later (in chapters 8 and 9) begin to discuss how these structures can be predicted, but for now it's important to understand that all molecules have a particular arrangement of their atoms and we will later learn how to predict and better understand the structures of many kinds of molecules.



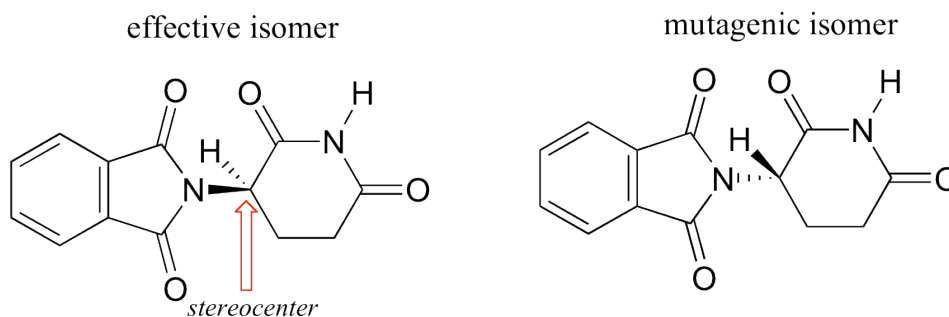
**Fig. 1:** Examples of molecular structure, a.  $\text{CO}_2$  (linear shape), b.  $\text{H}_2\text{O}$  (bent shape), c.  $\text{CH}_4$  (tetrahedral shape).

Some formulas, like  $\text{C}_2\text{H}_6\text{O}$ , can represent more than one stable molecule. Different molecules sharing the same molecular formula are called **isomers**. Ethanol and dimethyl ether are isomers with the formula  $\text{C}_2\text{H}_6\text{O}$  (Fig. 2). You might say they differ in that one is a poison and the other is not, though ethanol can be pretty poisonous too! Dimethyl ether would be harder to drink though because it is a gas at room temperature.



**Fig. 2:** Examples of isomers of  $\text{C}_2\text{H}_6\text{O}$ ; a. ethanol, b. dimethyl ether.

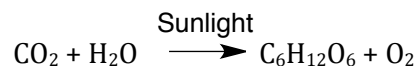
Isomers, especially ones with different types of arrangements of their atoms often have very different properties. Some isomers though have similar properties, sometimes even identical. One special form of isomerism are molecules that are non-superimposable mirror structures of each other, which are like molecular versions of a pair of hands. Like hands, these molecules have identical physical properties, such as melting points, boiling points, densities, etc., though they ultimately differ in reactivity when inside a chiral environment, or when they come in a contact with something like a “glove”, perhaps an enzyme with an active site that only fits one of the isomers. Examples are the mirror images of the molecule Thalidomide (Fig. 3). The effective stereoisomer was developed as an anticancer drug and later found use as a prenatal drug for pregnant woman to help with morning sickness. Unfortunately, the body serves as a catalyst to mutate the isomer into its mirror image, which ultimately proved very harmful to fetuses. Over 10,000 cases of severe birth defects were attributed to the drug. Thalidomide was only offered for a short period of time to pregnant women in the late 1950s, though the drug is still used today to fight some forms of cancer.



**Fig. 3:** Effective and mutagenic stereoisomers of thalidomide

Hopefully these example show that molecules have particular structures and that the properties of the molecule depend on its structure. In chapters 2-4, we will present you with some differences classes of compounds and describe how to name the compounds and we will discuss their bonding and molecular structures in chapters 8 and 9.

**Chemical reactions:** Chemical reactions describe the transformations of matter. We can often think of reactions that occur in nature or as reactions out that can be carried out in the laboratory. An example of a reaction that readily occurs in nature is the formation of sugar and  $O_2$  from the combination of  $CO_2$  and  $H_2O$  by photosynthesis through the absorption of sunlight by chlorophyll:



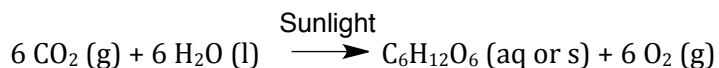
This reaction requires sunlight or else no products can be formed. The reaction is also not balanced, which means the number of atoms on both sides of the reaction are not equal. A reaction should describe the reactants that are initially present on the left side and the products that are formed by the breaking and remaking of chemical bonds. The reaction should have an equal number of atoms on both sides, since matter can neither be created nor destroyed in a chemical reaction.

To balance the reaction, coefficients can be added before each molecule to indicate the number of molecules required to put each atom into balance on both sides of the reaction. The above reaction requires six  $CO_2$  molecules to balance the

six C atoms in one molecule of  $\text{C}_6\text{H}_{12}\text{O}_6$ . Six  $\text{H}_2\text{O}$  molecules are needed to balance the twelve H atoms in  $\text{C}_6\text{H}_{12}\text{O}_6$ . And finally, six product  $\text{O}_2$  molecules balances the 12 O atoms in water and six O atoms in  $\text{C}_6\text{H}_{12}\text{O}_6$ .



Usually the **physical states** of substances in a reaction are indicated as either solid, liquid, gaseous, or aqueous (aqueous describes a substance dissolved in water). Sometimes these states can be inferred based on properties of the compounds and nature of the reactions. In photosynthesis,  $\text{CO}_2$  is absorbed by the plant from air, so  $\text{CO}_2$  is as a gas in the reaction. Water is pulled up from the roots of the plant, so it is a liquid in the reaction. The product sugar molecule can either be dissolved into water if enough water is present or it could a solid if most of the water has evaporated or reacted within the plant. Oxygen,  $\text{O}_2$ , is a gas at room temperature so it is a gas in the reaction.



**Tip:** If you are curious to know if a particular molecule exists as a solid, liquid, or gas at room temperature, try looking up the melting point and boiling point of the compound. You can look these up several ways: the CRC Handbook (best for inorganic compounds), the NIST Chemistry WebBook (best for organic compounds), and Wikipedia (a surprisingly great source for nearly all organic and inorganic compounds). The first two sources can be found through links in our course's library link in Carmen. A compound will be a solid at room temperature if its melting point is higher than room temperature, a liquid if its melting point is below room temperature and boiling point is above room temperature, and a gas if room temperature is above both its melting and boiling points.

There are three primary factors to consider that govern whether a reaction may or may not take place spontaneously:

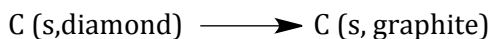
1. Are the bonds in the reactants or products more stable?
2. Are the reactants or products more disordered?
3. Is there an external energy source, such as electricity, light, or heat being applied to the reaction?

Most reactions that take place do so because the bonds of the products are more stable than the bonds in the reactants. Though, it is possible that an increase of disorder of products, which is thermodynamically favored because the act of increasing the order of system requires an increase in energy, can counter a reaction where the bonds of the products become less stable than in the reactants. Reactions that lead to the formation of less ordered products with less stable bonds, such as in photosynthesis, require an intense energy source as an external energy supply. The opposite type of reaction to photosynthesis is combustion. A combustion reaction leads to the release of heat, which can be observed by the flame of these reactions, such as the burning of butane in a lighter.

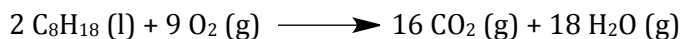


Whether a reaction *can* take place *spontaneously* often does not determine if it *will* take place at any appreciable *rate*. We usually say a reaction is *spontaneous* if it can occur based on the relative stability and randomness of the products compared to the reactants, or if energy is supplied to overcome the relative instability or ordered nature of the products. The *rate* at which a spontaneous reaction will take place though is mostly unrelated to its thermodynamics. A simple analogy is that the thermodynamics is like the total distance between two cities as it is the difference in energy between reactants and products. The rates of reactions though are most related to a topology map or the amount of energy required to rearrange the atoms of the reactants into products, similar to the height of a mountain that must be overcome upon traveling between two cities.

The rate of a reaction is most related to the topology of the chemical mechanism, or the pathway the reactants must take on route to the products. As an example, diamond and graphite are both allotropes of the element carbon (allotropes are different forms or structures of the same element). Both allotropes are solids, though diamond is of course much more precious. Graphite is just slightly more thermodynamically stable than diamond. Luckily though, a tremendous amount of energy is required to be absorbed by diamond to allow it change forms into the more stable form of graphite. So this reaction does not truly take place with an appreciable rate under normal temperature and pressure conditions.



Combustion reactions are an interesting type of reaction, since we certainly know these reactions release a tremendous amount of heat, which lead to their use as fuels for internal combustion engines. One of the more commonly combusted compounds is octane,  $\text{C}_8\text{H}_{18}$ , where the liquid can be burned in the presence of oxygen to yield carbon dioxide and water:



Despite the products having both a greater disorder and more stable bonds, octane will only violently combust with oxygen only if a small spark or amount of heat is applied.

In Chapter 3, we will discuss some other types of chemical reactions and some calculations that can be made on the products/reactants in a reaction (such as how much reactants must be mixed to form a certain amount of a product, how much product can be made from reaction of a given amount of reactants, etc.).

In chapter 4, we will classify and discuss some reactions that occur in aqueous solutions, such as acid-base reactions, metathesis (or precipitation-forming) reactions, and oxidation-reduction reactions.

In chapter 5, we will discuss heat changes that accompany chemical reactions. We will see that heat is released when more stable bonds are formed, though heat must be absorbed for a reaction to form less stable bonds. In Chapter 14, we will introduce reaction kinetics and see some examples of reaction mechanisms. In Chapter 19, we will discuss how disorder impacts a reaction and we will introduce a thermodynamic function (i.e. Gibbs energy) that directly predicts the spontaneity of a reaction.