Chemical reactions quiz answers

Questions about this quiz are based on information from Chemistry: Chemical Reaction. 10 Your Score Quiz Question: The correct answer: X ✔ Your answer: The correct answer: X ✔ Your answer: The correct answer: X ✔ The correct answer: X ✔ Your answer: The correct answer: X ✔ ✔ Your answer: Correct answer X ✔ Your Answer: The right answer: X ✔ Your answer: The correct answer: X ✔ Your answer: X ✔ Your answer: For a webquest or practice, print a copy of this quiz in Chemistry: Chemical Reactions webquest. About this quiz: All questions about this quiz are based on information available in Chemistry: Chemical Reactions. Back to Science for Kids For each of the following questions or statements, select the most appropriate response and click its letter: Congratulations – you have completed the #2-1 PRACTICE Quiz: Chemical Response Type. You scored %%SCORE%% of %%TOTAL%% (%%PERCENTAGE%%). Your performance was rated as: %%RATING%% Your answers are highlighted below. There are 10 questions to solve. If you look at this message, this means we're having trouble loading external resources on our website. If you are behind a web filter, please make sure the domain *.kastatic.org and *.kasandbox.org are not blocked. Think you have your head wrapped around Chemical Reactions? Put your knowledge to the test. Good luck – Stickman counts you! Reactions and Changes by Anthony Carpi, Ph.D., Adrian Dingle, B.Sc. Top Page 2 Reactions and Changes by Anthony Carpi, Ph.D., Adrian Dingle, B.Sc. (This is the latest version of the Chemical Reaction module. For previous versions, see this page.) The chemical reaction occurs altogether everywhere. Although we sometimes associate chemical reactions with the sterile environment of test tubes and laboratories - nothing could be further from the truth. In fact, the number of colossal transformations makes a whole variety of new material and the incomprehensible energy changes that occur in our world every second of the day. Naturally, chemical reactions can be less restrained than you can find in a laboratory, sometimes far away messier, and they usually happen whether or not you want them! Whether it be a raging fire throughout the forest (Fig. 1), the slow process of iron mourns in front of oxygen and water over a period of several years, or the delicate way in which the fruit cooks on a tree, the process of converting a set of chemicals (reactan) to another set of materials (products) is one known as a chemical reaction. Figure 1: A controlled fire in Alberta, Canada, is set to create obstacles for future fireflies. Image © Cameron Strandberg, Rocky Mountain House Although chemical reactions have occurred on Earth since the beginning it wasn't until the 18th century that early chemists began to understand them. Processes such as fermentation, in which chemical sugars are converted into alcohol, have been known for centuries; However, the chemical base of the reaction is not understood. What is this transformation and how are they controlled? These questions can only be answered when transitioning from alquality to chemistry as quantitative and experimental science occurs. Beginning in the early Middle Ages, European and Persian philosophers were blown away by the way some substances seemed transmute (or transformed) into others. Simple stones, such as those containing sulfur, seem magically burning; and otherwise, unimaginable minerals have changed, such as orange cinnamon becoming a dazzling silver liquid metal mercry when heated. Alchathists are based on their approach to Aristotle's ideas that everything in the world consists of four basic ingredients - air, earth, fire, and water (Fig. 2). Figure 2: Aristotle believes that everything in the world consists of four basic ingredients - air, earth, fire, and water. Therefore, they suggest, and spend a generation trying to prove, that less expensive metals such as copper and mercry can be turned into gold. Despite their perverse approach, many early alqualities perform basic chemical experiments - transforming one ingredient to another, and therefore it is difficult to point to a specific date or event as the birth of a ordered chemical reaction idea, can be quantified. However, there are some important moments in history that have helped make sense. Antoine Lavoisier was a French aristocracy in the 1700s who began experimenting with different chemical reactions. At the time, chemistry was still inevitable as true and quantitative science. Most of the theories exist to explain how the materials change depending on the Greek philosophy, and there is a detailed precious little experiment that attaches to alchemist tinkering. However, in the second half of the 18th century, lavoisier did many quantitative experiments and observed that although the material changed shaped during the chemical reaction, the mass of the system - or the measurement of the amount of goods present - had not changed. In doing so, Lavoisier championed the idea of mass conservation during transformation (Fig. 3). In other words, unlike alathymicists in front of him who think they created things out of nothing, Lavoisier suggested that the ingredients were not created or destroyed, but instead changed the shape during the reaction. Lavoisier's ideas were published in the seminal works of Traité élémentaire de Chimie in 1789 (Lavoisier, 1789), which was widely heralded as the birth of modern chemistry as science Rajah 3: Lavoisier Lavoisier Mass conservation, which states that the materials are not created or destroyed, but rather changed shape during reactions. In this example, the reactocant (zinc and two hydrogen chloride molecules) converts to different products (zinc chloride and dihydrogen), but no mass is lost or created. Joseph Proust is a French actor who follows in The Lavoisier's footsteps. Proust does dozens of chemical reactions, starting with different amounts of various substances. Over time, he observed that no matter how he started a certain chemical reaction, the ratio at which the reacments consumed always persist. For example, he worked extensively with carbonate brass and no matter how he changed the ratio of starting reakants, copper, carbon, and oxygen all responded together in a constant ratio (Proust, 1804). As a result, in recent years of the 18th century, Proust summarized the laws of continuous composition (also referred to as a definite proportional law, Figure 4). He realizes that any chemical given (which we now define as compounds) always consists of a similar ratio by the mass of parts of its elements regardless of the method of preparation. This is a big step forward in modern chemistry since it was previously believed that the materials formed during the chemical reaction were random and distraction. Figure 4: Examples of The Law Of Proust Continuous Composition, which states that any compound always consists of a similar ratio by the mass of parts of its elements, regardless of the method of preparation. English chemist John Dalton helped make sense of the mass conservation laws and definite proportions in 1803 by suggesting that the matter is made of unique material atoms that cannot be created or destroyed (see our module Preliminary Ideas on Things for more information). Dalton expanded Proust's idea by acknowledging that it was possible for two elements to form more than one compound, but whatever the compound, it would always contain elements combined in the overall number ratio (Dalton, 1808). This observation is known as the multi-proportional law (Figure 5) and with its atomic theory, helping cement the observations of the Lavoisier. Figure 5: The Multi-Part Dalton Law, which states that two elements can form more than one compound, but whatever the compound, it will always contain elements combined in the ratio of the overall number of this Progress, taken together, lay the ground for our modern understanding of chemical reactions, chemical similarities, and stoichiometry chemicals, or the process of expressing quantity in response and chemical products. Checkpoint understanding ____ was first the theory that although materials change shape during chemical responses, the system's variables do not change. There are a surprising variety of chemical reactions. Chemical Chemistry occur always in our bodies, in plants and animals, in the air circulating around us, in the lakes and oceans we swim, and even in the soil where we grow crops and build our homes. In fact, there are so many chemical reactions that happen that it will be difficult, if not impossible, to understand them all. However, one method that helps us understand them is to categorize the chemical reaction into some common type. Although it is not a perfect system, putting together reactions according to their similarities helps us identify patterns, which in turn allow forecasts to be made approximately as unentended reactions. In this module, we will consider and provide some context for some categories of response, specifically: synthesis, delegation, single replacement, double replacement, REDOX (including burning), and acid-based reactions. Regardless of the type of reaction, a universal truth applies to all chemical reactions. For the process is classified as a chemical reaction, that is, one in which chemical changes occur, new materials must be produced. The formation of new materials is almost always accompanied by energy changes, and often with some kind of physical or visible change. Physical changes can be of different types, such as the formation of gas bubbles, solid rain, or color changes. These changes are an indication of the existence of chemical reactions and are an important trigger for further research by chemists. Before lavoisier work, it is less understood that there are different gases consisting of different elements. On the other hand, various gases are usually misused as air types or parts that are missing air - for example, a commonly used term is inflammatory air, or dephlogisticated air. The lavoisier thinks differently and is convinced that this is a different substance. He conducted an experiment in which he mixed inflammatory air with dephlogisticated air and sparks, and he found that the ingredients combined to produce water. In response, he renamed the inflammatory air of hydrogen from Greek hydro to water and genes for creators. In doing so, Lavoisier has identified a synthesis reaction. In general, the synthesis reaction is one where the more easily combined material to form another more complex one. Hydrogen and oxygen (which Lavoisier also renamed dephlogisticated air) combine with the presence of sparks to form water, summarized by the chemical equations shown below (for more chemical similarities seeing parts called Chemical equation Anatomy), it represents a simple synthesis reaction. Equation 1 $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ in 1774, scientist Joseph Priestley turned curiosity over a mineral called cinnamon - the brick red mineral. When he put the mineral under sunlight strengthened by a powerful magnifying glass, he found that the gas produced which he described described has nature relegated due to candles baked in bright gas (Priestley, 1775). Without realizing it, Priestley has found oxygen as a result of the deployment reaction. Derogatory reactions are often thought to be the opposite of a synthesis reaction because it involves compounds broken down into simpler compounds or elements. In priestley oxygen case, he has broken the merzor (II) oxide (cinnamon) with heat into individual elements. Reactions can be summarized in the following equations. Equation 2 $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$ British chemist and meteorologist John Daniell, created one of the first practical batteries in 1836 (Figure 6). In his cell, Daniell uses a very common single replacement reaction. Its initial cells are a complicated affair, with unsustainable parts and complicated constructs, but on the other hand, the chemistry behind them is really quite simple. Figure 6: Daniell cell battery. In certain chemical reactions, a single constituent can replace another that already joins the chemical compound. Daniell cells work because zinc can replace copper in copper sulfate solutions, and thus perform electron exchanges used in battery cells. The reaction can be summarized as follows: Equation 3 $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$ This single shift is called a metal shift because it involves one metal replacing other metals, and many types of batteries are based on metal replacement reactions. However, some types of single replacement reactions exist, such as when metals can replace hydrogen from acid or from water, or halogens can replace another halogen in certain salt compounds. Controlled fire use is an important development for early civilization. Although it is difficult to lower the exact time that humans first recoact fire-producing burning reactions, recent research suggests it may have occurred at least a million years ago inside a South African cave (Berna et al. 2012). Chemically, burning is nothing more than a fuel reaction (wood, oil, gasoline, etc.) with oxygen. For burning there must be fuel and oxygen gas. However, this reaction often requires activation energy (discussed in more detail in the Chemical Bonding module: Chemical Bond Properties), which can be provided by 'spark' or energy sources for quailation. Fuel, oxygen, and energy are three things that make up what is known as fire triangles (Figure 7), and any one of them absent means that burning will not happen. Figure 7: The triangle consists of three things - fuel, oxygen, and energy. images © Gustavb In the modern world, a lot of fuel usually burned for energy, is hydrokarbon – a substance that contains hydrogen and carbon (as discussed in the detailed in our Carbon Chemistry module). Plants produce hydrocarbons when they grow, and thus make excellent sources of fire, and other hydrocarbons are produced when plants or animals resorate over time (such as native gases, oils, and other materials). When this fuel is combust, the hydrogen and carbon in it combine with oxygen to produce two very familiar compounds, water, and carbon dioxide. One easy example is the combustion of the original gas, or methane, CH_4: Equation 4 $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ Like burning all the ingredients of fire, heat and light are products, too, and they are products used to cook our food or to heat our homes. Each of the four types of responses above is a sub-category of one type of chemical response known as redox reaction. The redox reaction is one in which reduction and oxidation apply together, therefore the name. Individual processes of oxidation and reduction can be estimated in more ways than one, but whatever the definition, both processes are syndiotic, that is, they must take place together. In one definition, oxidation is described as the process by which a species loses electrons, and subtraction is the process by which a species gets electrons. In this way, we can see how the couple should happen together. If a chemical is losing electrons (and therefore being oxideed), then it must have other intersent chemicals that it can give people electrons. In that process, the second material (which gets electrons) is said to be reduced. Without such recognition of electrons, the original species must not lose electrons and no oxidation may take place. When electron reception is present, it will be reduced and the redox combination process completed. This type of Redox reaction can be summarized by a pair of equations - one to indicate the loss of electrons (oxidation), and the other to indicate an electron gain (reduction). Using daniell cell example above, Equation 5 Oxidation: $Zn \rightarrow Zn^{2+} + 2e^-$ Subtraction: $Cu^{2+} + 2e^- \rightarrow Cu$ Electrons indicated lost by zink in the first reaction, are the same electrons received by copper ions in the second. Together, the responses can be combined to cancel out electrons in both sides of the reaction, into the overall redox reaction: Equation 6 $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ Takrif other oxidation and subtraction also exists, but in each case, two parts of the redox reaction remain syndiotic - one loss and another advantage The loss of one species should not apply without the other species that gets. If soap will not easily produce lather in water, water is said to be 'hard'. Hard water causes all kinds of problems that just make it difficult to form a lather. The formation of compounds in water pipes (known as 'scale'), can block water flow and can cause in the industrial process. The textile manufacturing and beverage industry relies heavily on water. In such situations, the quality of water can make a difference to the final product, so controlling the composition of water is important. Hard water contains magnesium or calcium ions in the form of scattered salts such as magnesium chloride or calcium chloride. When soap (sodium stearate) comes into contact with one of these salts, it enters a multiple reaction that forms an insoluble rain known as 'soap scum'. Multiple reactions (also known as multiple replacement reactions) apply when two ionic ingredients come together and both materials be swap together. In general: Equation 7 $AB + CD \rightarrow AD + CB$ Where A and C are kasi (positively charged ions), and B and D are anions (subject to negative pay). In the case of soap response with calcium chloride, the response is: Equation 8 $CaCl_2(aq) + 2NA(C_{17}H_{35}COO)(aq) \rightarrow 2NaCl(aq) + Ca\ (C_{17}H_{35}COO)_2(s)$ Stearate calcium pepejal is what we call soap stim, which is formed by the reaction of sodium soluble salt (soap) in a double replacement reaction with calcium chloride. The basic reaction of acid takes place around, and also within us, all the time. From the classic low school baking soda volcano to the digestive process, we face acid and principles every day. When a hydrogen atom loses its only electron, it forms a positive ion, H^+. These hydrogen ions are an important component of all acids, and indeed one definition of acid is that of hydrogen ion donors. As well as scripted acids in lemon juice, etanoic acid in vinegar, or ordinary makmal acids such as hydrloric acid, they all provide their hydrogen ions in a chemical response known as acid-base retaliation. Contrary to chemical acids is known as the principle, and the principle can be estimated as the reception of hydrogen ions. Every time acid donates hydrogen ions to the base, the basic acid reaction has occurred, for example, when hydrochloonic acids donate hydrogen ions to bases such as sodium hydroxide: Equation 9a $HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$ Take a closer look at this reaction revealing that in water HCl provides H^+ as shown below: Equation 9b $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$ The resulting species, H_3O^+ (hydronium ions), may, on the contrary, it acts as an acid when in contact with any species that can receive hydrogen ions, such as hydroxide ions from sodium hydroxide: The equation 9c $H_3O^+(aq) + NaOH(aq) \rightarrow 2H_2O(l) + Na^+(aq)$ Combining the equations of #9a and #9b gives us the #9c equation. The #9c equation can be rewritten to indicate the individual ions contained in the solution, therefore: Equation 9d $H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow H_2O(l) + Na^+(aq) +$ Removing the audience ions from the equation above, we get a clean ionic equation: Equation 9e $H^+(aq) + OH^-(aq) \rightarrow \rightarrow$ Any chemical reaction that forms water from the reaction between acids and foundations as in the equation should #9e be known as neutralisation reaction. Checkpoint Understanding the types of chemical reactions in which a single constituent can replace another already participating in chemical compounds is: Chemical equations are always associated with chemical reactions because they are the acronym by which the chemical reaction is described. That fact alone makes the equation very important, but the equation also has an important role to play in describing the quantitative aspects of chemicals, something we officially call stoichiometry. All chemical reactions take the same and fundamental format. The initial material, or reactan, is listed using their chemical formula to the left side of the arrow, with multiple reactions separated by plus signs. In the case of a reaction between carbon and oxygen: On the right side of the arrow one finds a new material or material chemical formula (known as a product) produced by a chemical reaction. In this case, since carbon dioxide is the result of burning carbon with the presence of oxygen: The equation of 10b [Reactants] $C + O_2 \rightarrow CO_2$ [Product] Because the reaction can result in both physical as well as chemical changes, Each material is given a state symbol written as a subscription on the right side of the formula, this describes the physical form of the common state abbreviation is (s) for solids, (l) for liquids, (g) for gases and (aq) for any aqueous substance, that is, dissolved in water. Equation 10c $C(s) + O_2(g) \rightarrow CO_2(g)$ Finally, to ensure that these delegates comply with mass conservation laws, similarities may need to be offset by adding numbers in front of each species that creates the same number of atoms per element on each side of the equation. In the case of carbon dioxide formation from carbon and oxygen, there is no need for the addition of such numbers (called stoichiometric coefficient), since 1 carbon atom and 2 oxygen atoms appear on each side of the equation. Naturally, chemical reactions are often driven by exchanges in energy. In this regard, the response is generally separated into two categories - those that release energy and people who absorb energy. Exotic reactions are those that release energy into the environment (Figure 8, right). Burning reactions are a clear example because the energy emitted by the reaction is turned into light and heat seen in the immediate environment. On the other hand, endotermic reactions are those who absorb energy from the environment (Fig. 8, left). In these circumstances, one may need to heat the reaction or add some other form of energy to the system before viewing the outcome of the reaction. Figure 8: On the left is the endotermic response, where absorbed absorbed The environment. On the contrary, on the right is an extraordinary response, which releases energy into the environment. In both case it is important to note that energy is not created or destroyed, on the contrary it is transferred from one type of energy to another, e.g. from chemical energy to heat or light. The energy that goes into the formation of chemical bonds is exchanged for other types of energy with the environment around that reaction. A classic example is the act of photosynthesis, in which plants absorb light energy from the sun to realize the bonds between the atoms that make up the sugar, which are stored as chemical energy for later use by plants. The respiratory process is essentially the opposite of photosynthesis, in which the bonds in the sugar molecules break and the energy released is then used by plants. Comprehension Checkpoint ___ reactions are those that absorb energy from the environment. Chemical retaliatory acts are all around us every day. Whether it is a single replacement reaction in our light batteries, a synthesis reaction that occurs when iron rust is in front of water and oxygen, or a basic response to acid that occurs when we eat - we experience a chemical reaction in almost everything we do. Understanding these reactions is not an abstract concept for a chemist in a distant society, on the contrary it is important to understand life and the world around us. To truly master chemical reactions, we need to understand the quantitative aspects of this response, something called stoichiometrics, and the concepts we will discuss in other modules. This module explores various chemical reactions by collecting them into common types. We looked at synthesis, decomposition, single replacement, multiple replacement, REDOX (including combustion), and acid base reactions, with their respective examples. Main Concept Measures from qualitative science to quantitative, is important in understanding chemistry and chemical response more fully. If the material or material (reaktan), undergoing changes that result in the formation of new materials or materials (products), then the chemical response is said to have occurred. Jisim and energy are nurtured in chemical responses. It was not created or destroyed, instead it was preserved but reconnected to create new material. No energy is created or destroyed, it is preserved but often converted to different forms. Chemical responses can be classified to different types depending on their nature. Each type has its own pentakrifan characteristics in terms of reakmen and products. Chemical response is often accompanied by visible changes such as energy changes, discoloration, gas release or pepejal formation. Energy plays an important role in chemical reply. When energy is released into the environment said to be exothermic; when energy is inducted from the reaction environment is said to be endothermic HS-C5.4, HS-PS1. A2, HS-PS1. A3, HS-PS1. B3 Berna, F., Goldberg, P., Horwitz, L. K., Brink, J., Holt, S., Bamford, M., & Chazan, M. (2012). Microstratiography evidence in the fire there at the Acheulean Cave wonderwerk strata, North Cape province, South Africa. Proceedings of the National Academy of Sciences, 109(20), E1215-E1220. Dalton, John (1808). New Chemical Philosophy System. Lavoisier, Antoine (1789). Traité Élémentaire de Chimie, présenté and undre nouveau, et d'après des découvertes modern. Priest, Joseph (1775). Discovery Account Further in the Air. Philosophy Transactions. 65: 384–94. Proust Joseph Louis (1804). Sur les Oxydations Métalliques. J Phys. 59: 321-343. Anthony Carpi, Ph.D., Adrian Dingle, B.Sc. Visionlearning Chemical Reactions Vol. CHE-1 (6), 2003. Top Page 3 Chemical relations by Robin Marks, M.A., Anthony Carpi, Ph.D. Imagine that you are travelling in an exotic place and your car muffler dropped. You have to find a place where you can buy a replacement, and you don't speak local language. Forget the time now to paint a sketch that delivers that you're looking for a store where you can get a new section installed. Perhaps you include cars, mufflers, front shops, and someone holding a screwdriver or other tool. Then see your pictures: think about what type of understanding you are and the audience needs to share in order for you to convey your message. If you showed a similar picture to someone 500 years ago, they didn't know what the car was, let alone what a muffler did or that there were people who specialize in installing them. But you and your 21st-century audience can translate your sketches because of the shared knowledge of the car. Chemical equations play the same role for people who convey messages about what happens during the chemical response. You may bear in mind that In chemical reactions, bonds between atoms in compounds are broken down, and atoms rearrange to form new compounds, either release or deploy energy in the process (see our Chemical Response module). For example, let's consider what happens to the muffler. Just saying it's grounded there's not much explanation. You can say iron reacts with oxygen to produce rust. That's better, but not so accurate. What, exactly, is rust? Chemically, it is iron oxide, but iron forms many types of oxide. So we need a very specific way to express the chemical reactions that lead to the departure of our muffler. That's a chemical equation for: they're a kind of abbreviation used to deliver exactly what's going on in the reaction. In the most basic sense, the equation describe the type and amount of each material that responds (reaktan) to form a amount of certain ingredients produced (products). Reaktan will always respond in the civility given in the equation. If the supply of one reaktan runs out, the other overages will remain unable to be resented. The shorthand that describes the part of our rusty train is this: Large amounts in the presence of atoms or molecules, which we call multipliers, tell us the relative amount of each material involved in or produced by the response. The numbers in the subscription refer specifically to the elements that are before them. In other words, the 4 iron atoms in the muffler respond with 3 oxygen molecules in the air. Each of these oxygen molecules contains two oxygen atoms. When these chemical bonds reasociation and reaktan merge, the result is 2 specific iron oxide molecules containing 2 iron atoms and 3 oxygen atoms—a.k.a rust. Many of the irons in this muffler are combined with oxygen in the environment, manifesting rust and leaving holes in the muffler where the iron has been eaten. The © Raymond Webber's image of this equation conveys something much deeper than the number of zarah, though: it captures the centuries-old collection of knowledge of what our universe is made of and how it interacts. Like an elegant poem (or like your rusty muffler painting), the chemical equation conveys a world of complex concepts in just a few phrases. Point of Understanding Bush If one reakmen is actually eaten, the reaktan tray will As our understanding of the deep chemical processes over time, chemical equations slowly become more sophisticated. Ultimately, such equations play a role in the recognition of chemistry as its own important science, apart from medicine, alchemy (which was popular in the 17th and 18th centuries), and physical. The first known written chemical equation—which is really more than an image—appears in what is considered to be the first chemical textbook, Tyrocinium Chymicum (meaning The Beginning of Chemistry). It was written by French scientist and teacher Jean Beguin in Paris in 1615, and describes what Penguin noticed when he elongated the antimony sulfide with mercury chloride. Mercury becomes wap, leaving behind the remaining antimony oxychloride. Although the image of the Penguin stoning looks very different from modern chemical equations (and it is not entirely correct), it reflects an understanding of the cracks and products in chemical responses. Rajah 2: This painting, from Tyrocinium Chymicum Jean Beguin, shows what the author believes to be true when antimony sulfide responds with mercury chloride. Rajah's image conveys the understanding that chemical reactions begin with certain reactants and produce certain products. Beguin reveals the beginning of an understanding of what happened in the response but not an explanation as to why. More than a century passed before new diagrams revealed a deeper vision into the a phenomenon that drives reaction. The insight came from William Cullen, who was also a teacher, and founded the chemistry department at the University of Glasgow in 1747. Her hand-written lecture notes contain diagrams using arrows and letters, showing four different types of reactions: Figure 3:Drawing in a set of unknown lecture notes written by

William Cullen in the mid-18th century. Cullen developed a painting in the hope they could help his students better understand the chemical reaction. Image © University of Glasgow Library Diagram on the bottom left describes what Cullen called a single elective attraction, or what we now call a single replacement reaction, involving a single element that takes place of another element in the compound. (For more information about the types of reactions, see our Chemical Equations module. The top left is a diagram for a double replacement reaction, or that involves the exchange of components from each of the two different compounds. In both cases, the reaction produces precipitates, indicated by the squiggly line. Both diagrams on the right relate to the dissolution of salt. The current wisdom of Cullen's time is that elective attraction, which we now know as a charge, causes metal in one material attracted to other substances. Metals switch places, forming either new solids precious or dissolving in solutions (sometimes producing solutions that display new properties). While he doesn't have all the right details, Cullen's thoughts are on track, and his diagram represents an important step toward what is now a common generic chemical equation: We might take this kind of simple equation for granted now, but at the time, it shows some very looking ideas about what actually happened during the chemical reaction. Remember, atoms and molecules were still not understood in the days of Cullen. It takes three chemists working over the next 50 years to explain the pictures that give us modern chemical equations. So, to understand the idea that the ingredients combined and exchanged in a chemical reaction are really important. In 1774, French chemist Antoine Lavoisier made a crucial observation - she noted that although the substance in the chemical reaction changed in the way that Cullen described, the system's mass had not changed. In other words, the amount of each element present remains the same, which means that things and masses are preserved. This is an important concept that will be important when we discuss balancing chemical reactions - all elements should be fully accounted for at the beginning and at the end of the chemical reaction. Around the same time as Lavosier, Proust, another French man, works extensively with brass carbonate. She finds that regardless of how she changes the ratio of starting reacments - adding more copper sometimes, sometimes, More carbon or oxygen in others, copper, carbon, and oxygen all respond together in a constant ratio. His vision brings us a definite proportional law: in any given compound, the elements that occur in a fixed ratio, regardless of their resources. (Visit our module on Chemical Reactions to learn more about the work of this scientist). Again, this seems obvious to us today because we know that the elements only respond together in certain ways, but Proust completed his work before it was widely recognized that atoms and compounds existed. Finally, in 1803, English chemist John Dalton tied these threads together by suggesting that the matter was made of unique material atoms that could not be created or destroyed (see Our module preliminary Ideas on Things for more information). He points out that each element can combine with a number of others to form different compounds, and always in the ratio of the number of bongs. This knowledge, taken together, provides the basis for the chemical notations we use today. Chemical equations are unlike mathematical equations, which have been around for much longer. Although the quantity of each element must be the same on both sides of the equation, you will not see the same sign in the chemical equation. This is because the chemical equation explains the process of change. The Chemical Equation Checkpoint Understanding is written essentially in the same way today as they were in the 17th century. Like all other reactions, the creation of rusty muffler is an example of chemical change: Fe + O2 → Fe2O3 Reactants Product Revenue Again, iron and oxygen combine to form certain iron oxides. The arrow indicates that this reaction runs right as it is written, which means that iron oxide is formed. In some cases, the reaction can also backward, and we use a double arrow that indicates that some reactions go in both directions. Unfortunately for the muffler, that's not the case here. Now see the equation more closely: How much iron atoms on the side of the reaction (left) compared to the product side (right)? How many atoms of oxygen? You'll notice that they're not the same as the similarities that are being written. However, we know from the laws of conservation of things that atoms cannot be created or destroyed. In other words, we can't just get rid of iron atoms or make oxygen atoms to make common work. We also cannot change the subscription in the reaction, because doing so will suggest that we start with different reacments or get different products. All we can do is adjust the number of reactive components and (iron atoms and/or oxygen molecules on the left and iron oxide molecules on the right), because doing so does not envision realizing or destroying things. It is a way of ingyding - to produce molecules that have two iron atoms, you need to find the source of them Let's see how this works with the reaction that creates rust. We already know that the number of each type of atom is not the same on each side of the reaction. To deal with this, we can add multipliers in front of the reakmen and products to adjust the number of zarah and realize a balanced equation. (If there is no multiplier, this means there is only one of those types of zarah.) Let's look at the atomic reaction by atoms. Fe + O2 → Fe2O3 ↑ 1 iron ↑ 2 iron Since rust molecules have two iron atoms, we must compensate for Fe atoms by adding a 2-time multiplier in front of the iron atoms on the reaktan side, because that is the only place where iron appears to the left of the equation. Now we have two irons on each side of the equation. 2Fe + O2 → Fe2O3 ↑ 2 iron ↑ 2 iron Moving to oxygen atoms, we find two to the left of the equation, but three on the right. We can compensate for the equation by using one and a half oxygen molecules, as each molecule consists of two oxygen atoms. To better understand this, the image of one oxygen molecule as two atoms bound together O-O, Therefore, 1.5 oxygen molecules (O-O and O-O) provide three oxygen atoms: 2Fe + 1.5O2 → Fe2 O3 ↑ 3 oxygen ↑ 3 oxygen In the real world, though, oxygen does not apply in half the molecules. We can complete this conundrum by only familiarizing all the multipliers by two: 4Fe + 3O2 → 2Fe2O3 ↑ 4 iron ↑ 6 oxygen ↑ 4 iron & 6 oxygen Now we have 4 iron atoms on the left, and 4 (2 molecules, each containing 2 iron atoms) on the left, and 4 (2 molecules, each containing 2 iron atoms) on the left , and 4 (2 molecules, each containing 2 iron atoms) on the left, and 4 (2 molecules, each containing 2 iron atoms) on the left, and 4 (2 molecules, each containing 2 iron atoms) on the left, and 4 (2 molecules, each containing 2 iron atoms) on the left, and 4 (2 molecules , each containing 2 iron atoms) on the left, and 4 (2 molecules, each containing 2 irons For oxygen, there are 6 atoms on the left (3 molecules 2 atoms each) and 6 on the right (2 molecules containing 3 oxygen atoms). Now we have a balanced equation. Point of Understanding Bush To realize a balanced equation, we can together by telling us exactly how much chemistry is involved in the response, the balanced chemical equation tells us is the unconsciousness of the ingredients necessary to make a particular product. It's a bit like a recipe. Tell me you made a bunch of cookies. Recipe calling: 2 cups of flour 1 cup of sugar and promises you a bunch of 12 biskut. You can follow the recipe and use 2 cups of flour and 1 cup of sugar expect 12 cookies, or you can double the recipe and use 4 cups of flour and 2 cups of sugar and expect 24 biskut (or you can triple the recipe, or half, or more). Likewise, to create a group of 2 rust molecules, you need 4 iron atoms and 3 oxygen molecules: 4Fe + 3O2 → 2FeO3 ↑ 4 iron ↑ 3 oxygen molecules ^ 2 rust molecules So, for every four iron atoms and three oxygen molecules, we get two caramel molecules. As you would with a cookie recipe, you can multiply this: start with eight iron atoms and six oxygen molecules and create four molecular molecules Rust. No matter how many times you breed, the basic part is always constant. In the real chemical world, though, we don't deal with individual atoms and molecules; there are more than two rust molecules on mufflers that fall out of the car. That's where the concept of patronage is useful. Mole mole is a quantity of particles, especially lahidung 6.022 x 1023 particles. In other words, an iron atomic mole contains an atom 6.022 x 1023. Four iron atomic moles contain four times that amount, or 28,088 x 1023 iron atoms. (To brush your mole mathematics, see our module, Mole Mass and Atoms). Note that when we balance the equation, we think of the number of individual atoms on each side, attending the conservation laws of things that we cannot make or destroy the atoms. When we think about the proportion, however, we think of the circulation in front of the particles, which represent the number of moles of each type of particles used or produced. These cocales tell us how many moles the product can produce with the number of moles present at the beginning of the reaction. Chemical equations are a wise way to convey a lot of information in short order of parts. Modern chemical equations reflect our understanding of things that consist of atoms and chemical reactions as a process of breaking bonds and arranging atoms into new compounds. Mass is preserved in chemical reactions, and the number of particles on each side of the equation must reflect this. A balanced equation also delivers the proportion of the products and reactions involved in the reaction. Chemical equations are an efficient way to describe chemical reactions. This module describes the abbreviation notations used to express how atoms are modified to make new compounds during chemical reactions. It shows how balanced the chemical equation delivers the proportion of each reacment and the product involved. This module detects the development of chemical equations over the last four centuries as our understanding of the chemical process develops. Look at the chemical equations reveal that nothing is lost and nothing is gained in typical chemical reactions - things just change shapes. Robin Marks, M.A., Anthony Carpi, Ph.D. Visionlearning Chemistry Equation Vol. CHE-4 (7), 2018. Top Page 4 Chemical Relations by Anthony Carpi, Ph.D., Adrian Dingle, B.Sc. This is the latest version of our Chemical Bonding module. For previous versions, please go here. Life on Earth depends on water - we need water to drink, shower, cool ourselves on a summer day (Fig. 1). In fact, evidence suggests that life on Earth begins in water, more specifically in the ocean, which has a combination of water and salt, the most common table salt - sodium doing this usual water and salt appear on the organizing of great elements, periodic tables? Well, and millions of other substances, not found at its most famous of all chemical references: periodic tables. why not? The answer is a simple one. Figure 1: Life on Earth depends on water, not only for major biological functions but also for pleasure. For example, this relaxing oasis in the Mediterranean Sea, the cala Tío Ximo beach in Benidorm, Spain. Image © Table Diego Delso Table periodicly organizes 118 recognized chemical elements now, but water and sodium chloride are not elements. On the other hand, both are substances that consist of a combination of elements in a fixed ratio. The combination of a fixed ratio of 118 such elements is known as compounds. In its chemical reactions and physical interactions, sodium chloride does not act like the elements that make up it (sodium and chlorine); on the other hand, it acts as a completely different and unique substance. That's a good thing because chlorine is a poisonous gas that has been used as a chemical weapon, and sodium is a very reactive metal that is lightly explosive with water. So what allows sodium chloride to act in a completely different way? The answer is that in table salts, sodium and chlorine are accompanied together by a chemical bond that creates a unique compound, very different from the individual elements that consist of it. The chemical bond can be thought of as a force that holds atoms of various elements together in the compound. It opens up the possibility of millions and millions of combinations of elements, and the creation of millions and millions of new compounds. In short, the existence of chemical bonds contributes to chemical wealth that reaches far beyond just 118 building blocks. When discussing the history of chemistry, it is always dangerous to point to certain origins of the idea, since by its definition, the scientific process depends on the gradual refinement of the ideas that came before. However, as is the case with some such ideas, one can point to certain seminal moments, and in the case of chemical bonding, the early publications of the famous 18th century provided a time. In the 1704 Opticks publication, Sir Isaac Newton mentioned the force pointing to the modern idea of a chemical bond. In Inquiry 31 of the book, Newton described the 'force' - apart from magnetism and gravity - which allowed 'particles' to interact. In 1718, while translating the Opticks into his native language, French chemist Étienne François Geoffroy created the Affinity Schedule. In this striking first glimpse of the possibility of certain interactions, Geoffroy rounded up the relative affinity that various substances for other substances, and therefore reflect the strength between the materials. Current and Geoffroy's work predates our modern understanding of the elements and compounds, their work provides insight into the nature of chemical interaction. However, it is more than 100 years before the concept of combining the power of elements is understood in a more modern sense. In a paper in the journal Philosophical Transactions titled On a new series of organic bodies containing metal (Frankland, 1852), Edward Frankland describes combining the power of elements, a concept now known as a vacancy in chemistry. Frankland sums up his thoughts by suggesting what he describes as 'legal': Predisparency or law that happens (here), and that, no matter what atomic characters uniting might be, the power of combining interesting elements, if I can be allowed terms, is always satisfied with the same amount of atoms. Frankland's work suggests that each element combined with only a number of other elements atoms is limited, thus referring to the concept of bonding. But it is two other scientists who do the most important contemporary research on the concept of bonding. In 1916, U.S. scientist Gilbert N. Lewis published a paper now famous about the bond titled The atom and the molecule (Lewis, 1916). In the paper he outlined some important concepts about the bonds that are still being used today as a model of electron arrangement work at the atomic level. Most notably, Lewis developed a theory about bonding based on the number of outer shells, or valve, electrons in the atoms. He suggested that a chemical bond formed when two atoms shared a pair of electrons (later renamed a covalent bond by Irving Langmuir). Her Lewis dot diagram uses a pair of dots to represent each shared electron pair that forms a shared tie (Figure 2). Figure 2: Lewis dot structure for elements in the first two periods of the periodic schedule. The structure is written as a symbol of element surrounded by dots representing valve electrons. Lewis also championed the idea of 'octagonal' (group eight), that a valve shell filled was crucial in understanding electronic configuration as well as the way the atomic bond together. The oct was discussed earlier by chemists like John Newland, who felt it was important, but Lewis advanced the theory. Comprehension Checkpoint Lewis was based on his bond theory while still in college, a young chemist with the name of Linus Pauling familiarized himself with Lewis' work and began to consider how it might be interpreted in the context of a newly developed quantum mechanics field. Quantum mechanical theory, developed in the first half of the 20th century, has redefined our modern understanding of the atom and so any theory of bonding complete if it is inconsistent with this new theory (see our module Atomic Theory II:II: and The Beginning of Quantum Theory and Atomic Theory III: Duality of Waves and Electrons for more information). Pauling's greatest contribution to the field was his book The Nature of the Chemical Bond (Pauling, 1939). In it, he attributes quantum mechanical physics to the chemical properties of electron interactions that occurred when chemical bonds were created. Pauling's work is based on the establishment that real ionic bonds and peace bonds sit at the other end of the bond spectrum, and most chemical bonds are classed somewhere between the extremes. Pauling then develops a gelongsor scale of the type of bon that is administered by the electronegativiti atoms that take part in the bon. Pauling's great contribution to our modern understanding of chemical bonds led to him being awarded the 1954 Nobel Prize for his investigation into the nature of chemical bonds and his application to the judgment of complex material structures. Chemical bonds and interactions between atoms can be classified into several different types. For our purposes, we will focus attention on the same two types of chemical bonds, which are covalent and ionic bonds. Molecular bonds are formed when the juzuk atoms come close enough together that the outer electrons (injap) of an atom are attracted to the positive nuclear charge of its neighbour. As free atoms approach each other, there is both repulsive power (between the electrons in each atom and between the nuclei of each atom), and the power of attraction (between the positive nuclei and the negative injap electrons). Some juzuk require additional energy, called activation energy, to overcome the initial replanting power. But at various distances, atoms experience different interesting and recovered power, eventually finding an ideal separation distance at which electrostatic powers are reduced to a minimum. This minimum represents the most stable position, and the distance between atoms at this time is known as the length of the bon. As the name suggests, covalent bonds involve the sharing (co, meaning together) of injap electrons (outside the cangkerang). As explained earlier, the atoms involved in covalent bonds organize themselves to achieve the most powerful stability. And electrons are shared - sometimes the same, and sometimes not the same - between neighboring atoms. The easiest example of covalent bonds is when two hydrogen atoms come together to finally form a hydrogen molecule, H2 (Rajah 3). Rajah 3: Here the interaction of two gaseous hydrogen atoms is sought showing potential energy (purple line) compared to the internuclear distance of the atom (in pm, trilionths one meter). The minimum observed in power is potentially indicated as the length of the bon (r) between atoms. image © Saylor Covalent Academy in hydrogen molecules are defined by a valve electron pair (one from each hydrogen atom) shared between atoms, thus giving each atom of hydrogen a valve shell filled. Since one shared electron couple represents a bundle of chopsticks, hydrogen atoms in hydrogen molecules are held together with what is known as intermittent bonding, and that can be represented by a row, h-H. There are many instances where more than one pair of valve electrons are shared between atoms, and in this case some joint bonds are formed. For example, when four electrons are shared (two pairs), a bond is called a covalent double bond; in the case of six shared electrons (three pairs) of the bond is called a triple tie. A common example of such double bonds is the one that forms between atoms in oxygen and nitrogen gases. In oxygen gas (O2), two atoms share multiple bonds that produce structureS O=O. In nitrogen gas (N2), a three-tie exists between two nitrogen atoms, N≡N (Figure 4). Figure 4: Bonding between oxygen atoms and gas nitrogen. In oxygen gas (O2), two atoms share double bonds that produce structureS O=O. In nitrogen gas (N2), a three-tie exists between two nitrogen atoms, N≡N. Joint double bonds are shorter and stronger than comparable single-wheeled bonds, and in turn, triple bonds are shorter and stronger than double bonds - nitrogen gas, for example, not reacting easily because it is a very bound stable compound. Understanding Checkpoint When four electrons are shared between atoms, the ____ bond is formed. The ionic bond occurs when valve electrons are shared so unevenly that they spend more time around their new neighbors than their original nucleus. This type of bond is classically described as it happens when an atom interacts with each other to either lose or get electrons. Those atoms who have lost electrons earn net positive charges and are called cages, and those who have got electrons earn negative charges are clean and referred to as anion. The number of electrons obtained or lost by constituent atoms usually adheres to the octocate of the Lewis valve, or the principle of a valve shell filled. In reality despite the most classic examples of ionic bonding, such as sodium chloride bonding, contains the characteristics of a covalent bonding, or the sharing of outer shell electrons. Common misconception is the idea that elements tend to bind with other elements to achieve this octtet because they are 'stable' or, even worse, 'happy', and that's what the elements 'want'. The elements have no such feelings; on the other hand, the real reason for the formation of bonds should be considered in terms of stability resulting from positively charged nucleic electrostatic interactions with negatively charged electrons. Materials held together together ionic bonds (such as sodium chloride) can usually separate into true ions when acting by external forces, such as when they dissolve in water. Further, in solid form, individual atoms are not attracted cleanly to one individual neighbour, but instead they form a giant network attracted to each other by electrostatic interactions between each atomic nuclei and neighboring valve electrons. The attractiveness between neighboring atoms gives a solid ionic of a highly ordered structure known as ionic lattice, in which charged particles instead line with each other to create a rigid, highly-bound structure (Figure 5). Figure 5: Sodium crystal chloride, showing the rigid, very well organized structure. Solid ionic licensing structures convey certain properties common to ionic materials. These include: High melting and boiling eyes (due to the strong nature of ionic bonds across the turbidity). The inability to run electricity is in solid form when the ions are rigidly held in a fixed position in a lattice structure. Solid ionic is an insulation. However, ionic compounds are often capable of running electricity when softened or in a solution when free ions move. The ability to dissolve polyunses such as water, some of which charge nature leads to attractions to the ions charged in erectile dystic. Special solid ionic features are discussed in more detail in the Solid Real Estate module. Comprehension Checkpoint Atoms lost electrons and secured a net positive charge and called Lewis using the point to represent valence electrons. Figure picture dot Lewis (see Figure 1) is a quick and easy way to show individual atomic valve electron configurations where no bonds have yet been made. Dot diagrams can also be used to represent molecules formed when different species bond with each other. In the case of molecules, the dot is placed between two atoms to describe a conciliatory bond, in which two dots (a pair of shared electrons) mark a conciliatory bond. In the case of the hydrogen molecules discussed above, the two dots in the Lewis diagram represent an electron couple shared and hence a bond (Fig. 6). Figure 6: Two hydrogen atoms connected by a kovalent tie. This can be represented by two dots (left) or a single bar (right). If the bonding of ionics and the kovalent bond sits at the extreme end of the tie-up spectrum, how do we know wherever a particular compound sits on that spectrum? Pauling theory relies on the concept of electroegativity, and it is a difference in electronegativity between important atoms in determining at any bond may be placed on a bond-type sliding scale. Pauling Scale assign numbers between 0 and 4 to each chemical element. Bigger The higher the electroegativity and the greater the attractions available for electrons. Electromagtivity differences between the two species help identify types of bonds. Ionic bonds are where a big difference in electronegativity exists between two species of bonding. Large differences in electrometectivity usually occur when metal tie-ups to non-metallics, so bonds between them tend to be considered ionic. When the electromelivitis differences between the atoms that make up the chemical bond less, then the partnership is considered a major interaction, and the bond is considered to be coincidental. Although it does not mean absolute, some consider the boundary between ionic bonds and covalents existing when electroegativity differences are around 1.7 – less differences tend towards coconut, and greater differences tend to be towards ionics. Smaller differences in electromagtivity usually occur between elements that are both considered non-metalant, so most compounds consisting of two non-metal metal atoms are considered covalent. Checkpoint Understanding If there is a big difference in electronegativity between two different elements, the bond between them will Once the difference in electronegativity has been considered, and bonds have been determined as covalent, the story is not so over. Not all conciliation bonds are created equally. The only true bond, perfectly going to be one where the electroegativity differences between the two atoms in the tie are equal to zero. When this happens, each atom has the same attraction for electrons that make up the bond together, and therefore the electrons are properly shared. This usually occurs in diatomic molecules (two atoms) such as H2, N2, O2, and those halogenic compounds when the atoms in the tie are the same. However, most collaborative bonds occur between the elements where although electroegativity differences are lower than 1.7, it is not zero. In this case, the electrons are still considered shared, that is, the bond is still considered perplexed, but the partnership is not perfect. Most conciliating bonds are formed between different atoms of electronegativity, meaning that joint electrons are attracted to one atom in a tie-up more than the other. As a result, electrons tend to spend more time at one end of the tie than the other. This establishes what is known as dipole, literally meaning 'two pillars'. One end of the tie is quite positive (less attraction for electrons), and one end of the tie is quite negative (more attraction for electrons). If this difference in electron affinity exists across the molecule, then the molecule is said to be pole - which means that it will have two different ones, and vice versa, a partial charge on both ends. Water (H2O) excellent examples of polar molecules. Electron Electrons not shared either since hydrogen and oxygen have different electromeagtivity. This creates dipoles in each H-O bond, and these dipoles do not cancel each other, leaving the water molecular poles in total (Figure 7). (Read more about these bonds in our Liquid Real Estate module.) Figure 7: In panel A, the water molecule, H2O, is indicated by the unearthed sharing of electrons resulting in partial negative charges around the oxygen atom and partial positive charges around the hydrogen atom. In panel B, three H2O molecules interact well, forming a dipole-dipole interaction between partial charges. When the electrons in the bonds are shared perfectly, there is no disposal, and do not end the bonds carrying any partial charges. When no such whole charge exists, the molecule is said to be unfolding. An example of a molecule that is not pole is hydrogen, H2. In larger molecules with multiple bloody bonds, each bond will not have apole or policeed with multiple degrees of partial charge. When all these dishes are accounted for in three dimensions, the distribution of uneven charges caused by the dipoles can cancel, making the molecule non-polar. Alternatively, there may be partial electricity charges across the molecules, making them pole molecules. Examples of non-atomic atomic atomic poles are carbon dioxide. Electrons are not shared either across bones C=O because carbon and oxygen have different electronivity. This creates dipoles in each Tie C=O, but because this is in line with linear molecules, with oxygen atoms on both sides of the carbon atom, they cancel through symmetrics to leave non-pole carbon dioxide molecules (Figure 8). Figure 8: Electrons are not shared either across the Tie-up C=O in CO2 and therefore it contains two dipoles. Since these two dipoles go against each other across linear molecules, they cancel through sympetry to leave non-pole carbon dioxide molecules. Image © Molecule: FrankRamspott/iStockphoto We have limited our discussion to ionic and covalent bonding and the sliding scale of the types of bonds that exist between them. However, many types of interactions and other bonds between atoms exist, especially metal bonding (an attraction that holds metal atoms together in metal elements), and intermolecular force (interactions that exist between, rather than inside, molecules bound together). Each involves electrostatic interactions similar to those described in ionic and consensus bonds, but even the extension is far from the end of the bonding story. In 2014, researchers found the first experimental evidence for a new type of interaction between atoms that had been predicted in the 1980s (Fleming et al., 2014). Named vibration bond, the theory is mild elements (in this case, hydrogen isotope) swings or bounces two heavier atoms (in this case, bromine) and effectively hold larger atoms together. Donald Fleming, a chemist based at the University of British Columbia in Canada, described the new bond as like a Ping Pong ball chanting between two bowling balls. As the investigation continues, we can expect to understand interactions at the molecular level with increasing sophistication, and with it, a greater understanding of what we call chemical bonds. The millions of different chemistry that make up everything on Earth consists of 118 elements that bond together in different ways. This module explores the same two types of chemical bonds: covalent and ionic. This module extends chemical bonds on a gelongsor scale from pure covalent to pure ionic, depending on the difference in electronegativity of bonding atoms. Highlights from three centuries of scientific investigations into chemical bonds include the 'power' of Isaac Newton, the dot structure of Gilbert Lewis, and the use of Linus Pauling's principle of quantum mechanics. The Main Concept When the power holds atoms together long enough to create a stable and free entity, that power can be described as a chemical bond. The 118 known chemical elements interact with each other through chemical bonds, to create a unique new set of chemical and physical properties that are completely different from the elements that make up them. It helps to think of chemical bonds as being on a gelongsor scale, where at one point there is a pure covalent bond, and on the other there is a pure ionic bond. Most chemical bonds are located somewhere between the two extremes. When a chemical bond is formed between two elements, the difference in atomic electronegativiti determines where on the gelongsor scale the bon falls. The big difference in electronegativity favors ionic bonds, there is no difference in realizing non-polar covalent bonds, and the rather small difference leads to the formation of polar caged bonds. HS-C4.3, HS-C6.2, HS-PS1. A3, HS-PS1. 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Top Page 5 Theory and Structure of The Atom by Adrian Dingle, B.Sc., Anthony Carpi, Ph.D. This is the latest version of my Atomic Theory module. For previous versions, please go here. By the late 1800s, John Dalton's view of the atom as the smallest particle that make up all things has been held for about 100 years, but the idea will be challenged. Some scientists working on atomic models found that atoms are not the smallest particles that make up things, and different atomic parts have very different characteristics. English scientist Michael Faraday could be considered one of the greatest minds ever in the field of electrochemicals and electromagneticism. Somewhat paradoxically, all faraday pilot work was carried out before the discovery of basic particles that this electrical phenomenon depends on. However, one of Faraday's earliest experimental observations was an important precursor to the discovery of the first subatomic, electron particles. As early as the mid-17th century, scientists had experimented with glass tubes filled with what was known at that time as a rare air. Rare air is referred to a system where most gas atoms have been removed, but where the vacuum is incomplete. In 1838, Faraday stated that when passing the current through such a tube, an electric arc was observed. The arc begins at a negative plate (known as cathop) and travels through the tube to the other charged anod (Faraday, 1838). In his experiment, Faraday observed a flash that started part of the way down the tube, and travelled towards the anod. This leaves the area between the cathon and the onset of unilterred luminescence, and is later known as the dark space of Faraday (Figure 1). Faraday could not fully explain his observations, and it took some further developments in terms of tube technology, before a greater understanding emerged. Figure 1: Emission of light in low pressure tubes caused by electric current. As Faraday sees, the tube shows the dark space between the light around the cathtod (left, the negative charged) and the anod (right, prosecuted positively). Image © Andrejdam/Wikimedia Comprehension Checkpoint Faraday is famous for discovering and naming electrons. In 1857, German glass timer Heinrich Geissler, while working for fellow countryman and physicist Julius Plücker at the University of Bonn, improved the quality of the vacuum that could be achieved in those tubes. However, geissler tubes still contain enough gas atoms that when electricity Wandering in a tube, there was an interaction between the two, causing the tube to glow. In the mid-19th century, this Geissler tube was largely nothing more than curiosity, but interestingly, a curiosity that proved to be a neon lighting streak. Englishman William Crookes repeats an experiment similar to Faraday and Geissler, but this time with a 'new and improved' vacuum. The number of gas atoms (and therefore pressure) has been drastically reduced in Crookes tubes. This causes an interesting effect: Faraday's dark space is observed further down the tube, again extending from the cathos towards the anod. In addition to the extension of the dark space, fluorescence is observed on the glass behind the anod at the positive end of the tube. When further experiments revealed that the shadow of an object placed in a tube was thrown onto the glass behind the anod, German physicist Johann Hittorf suggested that the shadows must be created by something wandering in a straight line from the cathop to the anod. Another German physicist Eugen Goldstein, burned the rays of this invisible beam cathon. As discussed in our module The Initial Idea on Things, Dalton's atomic theory suggests that the atom is invisible, that is, that it is the smallest particle that makes up things, and that all things are based on that single unit. Experiments with cathed ray tubes dramatically changed that view when they led to the discovery of the first subatomic particles. J.J. Thomson was an English physicist working with a cathosine ray tube similar to that used by Crookes and others in the mid-19th century. Thomson's (Thomson, 1897) experiment went further than he was ahead of him and gave evidence of the natures of something that Hittorf had hanged. Thomson notes that the cathed rays have been dismantled by the magnetic field and that the dissonance is the same no matter what the source of the rays. This suggests that the rays are universal in their properties and that they have some magnetic charges. Thomson further indicates that cathed rays are imposed because they can be dismantled by the electrical field. He found that the rays were hammered towards a positive plate and away from negative plate, thus determining that they consist of such negatively charged particles. Finally, Thomson uses both electrical and magnetic fields to cathed rays at the same time. Knowing the strength of the field used and measuring the auction of particle flow in a tube, he first was able to measure the velocity of the particles in the stream. Then, by measuring the ray dissonance while changing both fields, Thomson was able to measure the massive ratio to impose particles on the stream and he found something Negative particles have a massive ratio to charge the 1,000 times lower than the hydrogen atom, suggesting that the particles are very small – much smaller than the known sparked atoms. This fact allows Thomson to definitively say that atoms are not the basic building blocks of things and smaller (subatomic) particles exist. Thomson was originally called these particles of corpuscles, but later they were known as electrons. Comprehension Checkpoint J.J. Thomson determined that cathode rays consist of thomson's findings make sense of all previous observations made by Faraday, Geissler, and Crookes. Zipping through a tube filled with gas but partly under the vacuum, electrons will eventually slam into that gas atom, knocking down some of their electrons and making them fluoresce. The dark space observed by Faraday is due to the distance required for electrons to accelerate the speed required to ionise the atoms of tube gas. In a better vacuum achieved in the Crookes tube, electrons can travel further distances without interacting with gas molecules due to lower molecular density in the tube, thus prolonging the dark space. Less Gas Understanding Checkpoints are available in cathos ray tubes, dark space __ can be observed. With electrons now discovered, Thomson proposes a completely new atomic model known as the Plum Plum Ping Model. The model is called since it mimics the British desert of the same name that has dried fruits (especially non-plum raisins), dispersing in the bodies of sacks and eggs that make the dough. In its model Thomson suggests that negatively charged electrons (analogy to raisins) have spread randomly among so-called uniform positive electrical spheres (analogy with dough or pledging body) (see Figure 2). Figure 2: the atomic Thomson plum plum plum model, showing a positively charged sphere contains many negatively charged electrons in random order. The Thompson model of atoms as a dough cluster of positive and negative particles continued until 1911, when Ernest Rutherford, a former Student Thomson's, advanced atomic theory yet notched. In 1908-1911, Ernest Marsden and Hans Geiger did a series of experiments under the direction of Ernest Rutherford at the University of Manchester in England. In this experiment the alpha particles (small, positively charged particles) are fired on a thin piece of golden foil (Fig. 3). Under the Thomson Plum Plum Plum atomic model, the uniform positive electric sphere is considered so diffuse that alpha particles move small, moving fast going through. Similarly, electrons in this model are considered so small that any electrostatic interaction between them and positive alpha particles will be minimal, so the route alpha particles will not be affected. Figure 3: Goldfly experiments designed by Rutherford, Marsden, and Geiger. A positively charged alpha particle beam is shot on a piece of gold foil. The screen around the foil captures the effects of alpha particles. As predicted, Rutherford and his colleagues observed that most of the alpha particles were passed straight through the gold foil, and some particles were hammered at a small angle. However, contradicts what the Plum Plum Plum model predicts, some bounce back at a very sharp angle, some even fly straight towards the source! These particles act as if they face hard objects, such as tennis balls bouncing off the brick wall (Fig. 4). Figure 4: In the golden foil experiment, Rutherford and his colleagues are expected to see alpha particles passing through the most empty atoms of Plum Ping. However, what they note is that alpha particles are sometimes nailed at sharp angles, indicating there is something firmer in the atom than previously thought. The fact that most alpha particles passed directly through gold foil suggests to Rutherford that atoms consist of largely free space. However, contrary to the Plum Thomson Pding model, Rutherford's work suggests that there are dense, positive areas imposed in atoms that cause insulation and backscattering of alpha particles. Rutherford was surprised by this observation and famously said: It was the most incredible event that ever happened to me in my life. It's almost unusual as if you're shooting a 15-inch shell on a piece of tissue paper and it comes back and hits you. On consideration, I realized that this scattered backwards must be the result of a single collision, and when I made the calculation I saw that it was impossible to get any order of magnitude unless you take a system where a larger part of the atomic mass is concentrated within a minute of nucleus. At the time, I had an atomic idea with a massive central minute, bringing charges. More than a series of experiments and papers (Rutherford, 1911, 1913, 1914), Rutherford developed an atomic model with a dense, positive atomic area in the middle, now known as nukleus - and the atomic nuclear model was born. Comprehension Checkpoint Rutherford and his colleagues were surprised that Following the discovery of electrons, Nobel Prize-winning physicist Robert Millikan conducted a clever experiment that allowed certain values of electron negative charges to be calculated. In a well-known oil drop experiment, Millikan and colleagues sprayed small oil drops from the atomizer into a sealed room (Millikan, 1913). Oil drops fall down, under the influence of gravity, into the space between two electric plates. Electrical. they were charged, by interacting with the air that had been ionted by X-rays. Figure 5: Millikan oil drop experiment where he observes oil drops falling between two electric plates, where the drops become ionized by X-rays. By adjusting the voltage between the two electric plates, Millikan uses an electric force upward that exactly matches the gravity force down, thereof hanging the drops without movement. When suspended, electric power and gravity force work in the opposite direction but equally in magnitude. Therefore: where q is a charge on the decline in oil, E is an electricity field, m is the mass of the decrease in oil, and g is a field of gravity. By measuring the mass of each oil decrease and knowing both gravity and electricity fields, charges at each decrease can be determined. Millikan found there are different charges on different oil drops. However, in each case, charges on oil drops were found to multiples of 1.60 x 10-19 coulombs. He concluded that different charges were due to different numbers of electrons, each of which had a negative charge of 1.60 x 10-19 coulombs, and thus charges over electrons were found. Comprehension Checkpoint Millikan found that a different drop in oil in experimental Thomson electrons and Rutherford's nuclear model was a tremendous progress. Japanese scientist Melo Nagaoka previously dismissed model Plum Thomson's Plum Pding on the grounds that resisting allegations could not penetrate each other, and he proposed an atomic model that resembled the planet Saturn with an electron ring revolving around a positive center. After listening to Rutherford's work, he wrote to him in 1911 saying, Congratulations on the simplicity of the appdas you took and the outstanding results you got. But, the planet's models aren't perfect, and some inconsistent experimental observations mean a lot of work is still done. By the time the electrons are still thought to be small particles, and they are considered to rotate almost randomly around the atomic nucleus. It will take additional experiments and genius Neils Bohr, Max Planck, and others to make paradima shifts from classic physics where atoms consist of small particles and are administered by motion laws, quantum mechanics where electrons behave like waves and exbits (See our interactive animation comparing orbital and quantum models of the first 12 elements.) To learn more about the strange behavior of quantum physics, read other entries in our Atomic Theory series: II: Bohr and The Onset of Quantum Theory, III: Duality Waves-Particles and Electrons, and IV: Quantum and Orbital Numbers. Interactive Animation: The atomic and ionic structures of the first 12 elements of the 19th and early 20th centuries saw advances in our understanding of atoms. This module took the reader through experiments with cathode ray tubes that led to the discovery of the first subatomic zarah: electrons. The module then depicts the Thomson atomic plum pudding model alongside Rutherford's golden experiments which resulted in an atomic nuclear model. Also described is the Millikan oil decline experiment, which allowed it to determine electron charges. Readers will see how the work of many scientists is critical in this period of rapid development in atomic theory. The Main Atomic Concept is not a dense spherical but consists of smaller particles including negatively charged electrons. Research on passing electric currents through vacuum tubes by Faraday, Geissler, Crookes, and others lays the ground for the discovery of the first subatomic zarah. J.J. Thomson's observation of cathode rays provides the basis for electron discovery. Rutherford, Geiger, and Marsden conducted a series of golden experiments showing that atoms have small, dense, positive centers - then called nucleuses. Millikan oil reduction experiments determined the base charge on electrons as 1.60 x 10-19 coulombs. HS-C4.4, HS-C6.2, HS-PS1. A1, HS-PS1. A3 Faraday, M. (1838). VIII. Research experiments in electricity. The 13th series. Royal Society of London Philosophical Transactions, 128: 125-168. Millikan, R.A. (1913). On basic electrical charges and Avogadro Pemalar. Fizik Reviews, 2(2): 109-143. Rutherford, E. (1911). Spreading α β the particles by the matter and the structure of atoms. Philosophy Magazine, Siri 6, 21(125): 669-688. Rutherford, E., &; Nuttal, J.M. (1913). Spreading the α by gas. Philosophy Magazine, Siri 6, 26(154): 702-712. Rutherford, E. (1914). Atomic structure. Philosophy Magazine. Series 6, 27(159): 488-498. Thomson, J.J. (1897). Cathode ray. Philosophy Magazine, Siri 5, 44(269): 293-316. Adrian Dingle, B.Sc., Anthony Carpi, Ph.D. Atomic Theory I Visionlearning Vol. CHE-1 (2), 2003. Top Page 6 Atomic theory and Structure by Adrian Dingle, B.Sc., Anthony Carpi, Ph.D. The earliest idea of things at the atomic level has been built over the centuries. Beginning with the ancient Greeks, and moving into the early 19th century, the story unfolds relatively slowly. (You can read more about this is in our module The Initial Idea of the Matter: From Democritus to Dalton and Atomic Theory I: The Early Days.) Although the levels are slow, it is important to understand that the process is a method as each scientist is built on previous ideas. This is gradually, a logic progress, in which the structure of atoms develops rather than becoming an easy idea, philosophically, through the ultra-sophisticated world of the Higgs boson found in the early part of the 21st century, representing a wonderful example of the evolution of scientific ideas, and the use of scientific processes. In fact, one can argue that the history, struggle, and achievements that are downgraded through the development of understanding things at the atomic level are quintessential stories of scientific methods. The story of the first atomic theory faces reproducable, scientific (evidence-based) evidence in the late 18th century. French chemists Antoine Lavoisier and Joseph Proust, with the Mass Conservation Law respectively in 1789 and The Sure Proportional Law in 1799, each laid the groundwork for John Dalton's English work on double proportional Law (Dalton, 1803). Given that many centuries have passed between the earliest ideas of atomic work and Dalton, it is fair to say that the evolution of atomic theory has become one gradually, with developments in a stable field rather than astonishing. But that will all change, and quite dramatically. The most intense period of progress took place between the late 19th and early 20th centuries, and it depended heavily on the work of a Danish physicist named Niels Bohr. Like many ahead of him, Bohr was built on the work of his predecessor, and for Bohr, part of that foundation was built by Ernest Rutherford. Based on a series of experiments, Rutherford proposes an atomic planet model in which electrons swirl around hard, compact nuleus (see Atomic Theory I: Early Days). Although the Rutherford model explained a lot of observations precisely, it was found to have disadvantages. Rutherford's planetary model of the atom is based on classic physics - a system dealing with physical particles, force, and momentum. Unfortunately, the same system predicts that orbiting electrons in a way that Rutherford described will lose energy, give radiation, and eventually crash into nucleus and destroy the atoms. For the most part, however, stable atoms, lasted billions of years. Furthermore, the radiation predicted by the Rutherford model will be a constant spectrum of every color - essentially a white light that when through prism will display all the colors of the rainbow (Fig. 1). Figure 1: When passing through prism, white light displays a spectrum of colors. But when pure gases of different elements are excited about electricity, since they will become when placed in a newly discovered electrical emissions tube, they emit radiation at different frequencies. In other words, different elements do not emit white light, they emit light of different colors, and when the light is passed through the prisma it does not produce a continuous rainbow of color, but the pattern of the colored line, now referred to as the spectra line (Fig. 2). Clearly, models tidak tidak with all observations, and Bohr made his business to address this inconserence. Figure 2: The visible light spectrum displayed at the top and the spectra line for three elements - hydrogen, neon, and iron - is below. Image © of the Neon Spectrum: Deo Favente In 1911, Niels Bohr (Figure 3) had just graduated from the doctorate at Copenhagen University and was invited to continue his work at the University of Manchester in England by Rutherford. Rutherford has advanced the atomic theory with a groundbreaking gold inflammatory experiment, but Genius Bohr is taking the Rutherford model and advocateing it again. Figure 3: Niels Bohr It's not Bohr who has come up with the original idea of an atomic planet model, but he is the one who takes the basic concept and uses new ideas about quantum theory to him. The jump is necessary to explain the new evidence that challenges the old model, and then summarizes a 'better' new model. Interactive Animation: Checkpoint Atomic Understanding Bohr Model atomic planet is based It's easy to think of light, and other forms of energy, as persistent. Turn on the dimmer switch on your lights, and the lights become gradually brighter. However, in the late 1800s, physicists began to suspect that this was not actually true. Classic physics models fail to accurately predict black body radiation; In other words, classical physics does not accurately predict the energy given by the object when it is heated. German physicist Max Planck solved the problem in 1903 by suggesting that black body radiation energy had to be quantified, that is, that it could only be released or inflamed in certain 'packets' linked to certain frequencies. This solves black body problems and is consistent with observed experimental data. Therefore, quantum mechanics are born. Despite the progress made by him and others using this idea, interestingly, Planck remains quite skeptical of the quantity energy over the years. He stressed that the calculations he has done, and the conclusion he has achieved, are somehow sophisticated mathematical tricks and ultimately old, classic models will prevail. After all, it has been around for about 200 years and has stood up to some pretty intense scrutiny. In 1905, Albert Einstein published a series of papers suggesting that light also exhibited quantum behavior (Einstein, 1905). Sometimes described as Einstein's mirabilis Annus (miracle year), paperwork taken together, and combined with Planck's work, allowed Bohr to marry atomic properties with physics to usher in new dawns of understanding in atomic theory. In 1913, building theories Planck and Einstein, Bohr suggested that the electrons themselves had been quantified - that it was does not exist anywhere around the atom (as suggested by rutherford models) but rather only available in certain positions, with certain energeticity. Electrons can switch to different positions, but only in screwed, defined measures. It cannot rotate in any location around the atomic nucleus but rather limited to certain areas of space - just as the planet in our solar system is limited to certain routes. As negatively charged, electrons are attracted to positive protons in atomic nucleus, and will typically occupy orbit, or pathways, in the atoms closest to the nucleus if they are available. The state, which has potentially low energy, is called soil conditions. By exposing electrons to external energy sources such as electricity emissions, it is possible to promote electrons from their soil conditions to other posts with higher potential energy, called excited states. These 'excited' electrons quickly return to a lower energy position (to regain stability associated with lower energy), and in doing so they release energy in certain frequencies that correspond to the energy difference between electron orbits, or shells (see quantum behavior simulation). Bohr's mathematical equation further predicts that electrons will not crash into the nukleus in a classic physics manner - and rutherford models - have predicted. This is another important awareness that makes the jump from one paradigma (classical physics) to a new one (quantum physics). Interactive Animation: The atomic structure and ionic discovery of Bohr's first 12 elements that Planck's quantum theory can be used for classic Atomic Rutherford models, and can take into account the lack observed in the original model, is another wonderful example of how scientific theory uses previous evidence, plus new experimental observations, to adapt, develop, and change models Science is usually developed by contemporary scientists building upfront work and, as Isaac Newton put it in a 1676 letter to Robert Hooke (both leading scientists of their time), by standing on the shoulders of the giants. Bohr's work built on their theories in front of him, and expanded them to explain the experimentally observed atomic line spectra in sensible mathematical evidence. Checkpoint Understanding In atoms, soil conditions are orbital with potential energy __ Although Bohr's work seems to explain the phenomenon of wonder-some line spectra, some lines have been observed in hydrogen spectra that doesn't fit Bohr's theory. At first glance this appears poked a hole in bohr's idea, but Bohr quickly offered an explanation. He suggested that the lines in the spectrum to which should not be taken into kira is actually caused not by hydrogen atoms, but by completely different species altogether. So, what are these different species, and how do they come? Nearly 30 years before Bohr published his famous trilogy of papers in the Philosophy Magazine and Journal of Science in 1913, the idea of zarah-zarah who had the ability to bring some kind of accusation had been established by Swedish scientist Svante Arrhenius and The English Michael Faraday. The worn zarahs have been collected ions. Atoms are electrically neutral, which means that the number of positive protons in any atom given together with the same number of negative electrons. Because positive charges are precisely canceled by negatives, atoms do not have an overall electrical charges. However, just as it is possible to transport electrons into higher orbits, as Bohr's work has shown, it is also possible to provide electrons with enough energy to fully overcome the nucleus pull and to remove them from the atom completely. This has the effect of offsetting electrical charges and resulting in the formation of species with overall positive charges - called simen. For example, sodium atoms may lose electrons to form a positively charged sodium kasi (Equation 1); the energy associated with the first electron taunt of any atom is called the first ionizing power. Na(s) → Na+(s) + e-Cations formed by mocking one electron can be ionized again by losing additional electrons, and in the process forming another ion, this time with a charge of 2+. The energy required for the second electron taunt is known as the second ionizing energy. Although it rarely applies to larger atoms (or indeed with the examples of sodium given above), in theory it is possible to eliminate all electrons from any given atom, which leads to a third, fourth, fifth, and other powerful ionization for atoms with a large number of electrons. Rajah 4: Uses hydrogen elements, e.g. kasi and anion. image © Jkwchui Once an electron (or electron) has been rejected from an atom in this way, it is acceptable to another atom, and thus, it can be moved from one atom to another. Just like releasing unbalanced electrons, receiving electrons causes atoms to become unbalanced in terms of their charges as well, and once again ions are formed. This time, ions have negative charges (more electrons than protons), and the species is called anion. (Kasi hydrogen and anion are shown in Rajah 4.) For example, neutral chlorine atoms, with the same number of protons and electrons, can receive electrons from external sources to form negatively charged chloride anions (Equation 2). + e- → Cl-(g) Ions have different wild properties when compared to atoms, and electron transfer from one atom to another is an example of how small changes in the atomic structure can make a huge difference in the behavior and properties of particles. For example, sodium metal is unstable, reacting violently with water and eroding together in the air. Therefore, coming into contact with sodium free metal will be very dangerous. Similarly, chlorine exists as a gas under ambient conditions and it is highly poisonous, the lung scars of anyone who breathes it (in fact, free chlorine gas is used as a chemical weapon in World War I). However, when both substances react to each other, sodium loss electrons form a cage, and chlorine receives the same electrons to form an anion. The two ions resulted then bonded together as a result of their allegations, and together created a very common substance - table salt, which is not reactive or poisonous. Comprehension Checkpoint __ has a positive charge, while ____ has a negative charge. In Bohr's work, the ionizing of one particular element, helium, proved to be the key to unlocking the explanation of the unexpected lines he observed in the hydrogen spectrum. When the helium atom, which has two electrons, loses electrons to form helium ions, He+, its electronic structure mimics atomic hydrogen, since both species have only one electron - helium ions are said to be isoelectronic with hydrogen atoms. However, the helium ion has a nueleus with twice the charge of hydrogen atoms (two protons versus one proton). Bohr realizes this, and suggests that greater attraction between electrons and He nucluses takes into account the spectrum lines that were previously unexplained - a nuctleus charge affects the energy associated with the electron transfer between orbitals. Bohr's theory proved correct when the spectra was produced using ionized helium that had offended hydrogen. At the close of the 19th century, two different particles were known to exist in the atoms, and both had electricity charges - very small and negatively charged electrons and far larger and positively charged protons. However, in the early 20th century, evidence began to mount that this was not a complete overview of the atoms. In particular, the mass of protons and electrons in the atoms does not seem sufficient to justify the overall mass of the atom, and some kind of nuclear damage suggests that something else may occur in the nucleus. In 1932, James Chadwick, a British physicist who had learned with, and worked for, Ernest Rutherford at the time, set out to solve the problem. Rutherford had proposed the idea of a neutral atomic particle that had a mass as early as 1920, but he never had gaining traction in the hunt for this mystery. In 1932 1932 developing experiments that were first performed by Frederic Joliot-Curie and Irene Joliot-Curie. They found that by using polonium as a source of alpha particles, they could cause beryllium to remove radiation that, in turn, could be used to knock the proton out of a piece of paraffin candles. The Joliot-Curies suggest that this radiation is gamma radiation, a packet of energy without true mass. As researchers accomplished gamma rays and atomic nuleus, Chadwick realized what others didn't - that protons were too large to be pushed from paraffin by less massive gamma rays. By more carefully measuring the effects of mysterious particles on paraffinic candles and combining this with other measurements, Chadwick concludes that the particles removed are not gamma radiation, but relatively heavy particles that have no charge - a particle named neutron (Figure 5). Figure 5: An artistic model of the atomic showing the nuceleus, with protons and neutrons, and electron orbiting. Image © Visionlearning Chadwick wrote a paper about its findings titled Neutron's Possible Existence, and it was published in the journal Nature (Chadwick, 1932). In 1935, he was awarded the Nobel Prize in Physics for his breakthrough. Joliot-Curies didn't go without recognition either. Their work on radioactive and radioactive isotopes won those Nobel Prizes for Chemistry, also in 1935. Chadwick's findings mark the incidence of a induced nuclear response, in which neutrons are accelerated and crashed into the nucleus of other elements, generating large amounts of energy (the neutron can easily do this because neutrals it is not expelled from the nucleus in a positive way imposing particles will). These reactions have had a major impact on the world as a whole since they shifted the thought of atomic bombs and nuclear energy. Comprehension Checkpoint A ____ is a heavy particle that has no charge. Neutron also explained the existence of the atoms of the same elements that have different atomic mass. Isotopes are the same elemental atoms (that is, they have the same number of protons) but that differs in the number of neutrons they have. As a result, different isotopes have the same chemical properties, but their crowd, and in some cases their physical behavior, is different. Isotopes are expressed by their atomic mass, which can be shown by writing symbols of elements, followed by a dash and then mass, or, more commonly, by writing mass as a superscript before a symbol of element. For example, carbon-12 (C-12, or 12C) and carbon-14 (C-14, or 14C) are both naturally occurring carbon isotopes. Carbon-12 is a stable isotope that accounts for nearly 99% of the naturally occurring carbon. Carbon-14 is a radioactive isotope that only about 1 x 10-10% of the carbon occurs naturally, but it decays to nitrogen with a half-life of about 5,730 years, it can be used to date some objects containing carbon. (Three carbon isotopes are pictured in Figure 6.) Figure 6: Carbon Isotopes. Each has the same number of protons, but the number of neutrons is different. Often there is more than one naturally occurring isotope of any given, individual element. As a result, the atomic mass given on a modern periodic table is the weighted mass of all known isotopes of each element. For example, the natural chlorine that occurs has two main isotopes, one with 18 neutrons (mass 35), and one with 20 neutrons (mass 37). The lighter Isotope 35Cl, accounting for nearly 76% of many natural, while the 37Cl isotope only accounts for about 24%. Therefore, the weighted meaning is the average mass of natural chlorine atoms that occur that are developed by many of their relatives, or 35.45. Comprehension Checkpoint Isotopes is the same element atom that has different Bohr work providing a bridge between a number of different ideas that may never be associated together without its intervention. He converted the paradigma from using classical physics (particles) to atomic models, to think of the use of quantum theories and waves - a truly important development in the great scheme of atomic theory and that lay the ground for future scientists to build. Having travelled from the earliest work of atomic theory through the early part of the 20th century, we still had some way to go in atomic tales. The progress described in this module is eroded in Rutherford's work, modified by Planck's vision, and surrounded together by the Bohr genius. Still, there will be at least another decade who had to pass before work by Pauli, Heisenberg, and eventually Schrödinger led to the full development of modern quantum mechanics depicting atoms as we know it today. This module is the latest version of our previous content, to see longer modules please go to this link. The 20th century brought a huge shift in our understanding of atoms, from planetary models that Ernest Rutherford proposed to edify quantum theories and Niels Bohr waves to electron behavior. With a focus on Bohr's work, the developments explored in this module are based on the progress of many scientists over time and lay the groundwork for future scientists to build further. The module also describes James Chadwick's findings from neutron. Among other topics are anion, cage, and isotopes. Main Concept Painting experimental and theoretical evidence, Niels Bohr transformed the paradigraph of modern atomic theory from one based on classic physical particles and physics, to one based on quantum principles. Under atomic models electrons cannot be free around the atoms, but is tied to a certain atomic orbit that both constraints and define atomic electronic behavior. Atoms can gain or lose electrons to become electrically charged ions. James Chadwick completes an early picture of the atom with the discovery of neutrons, neutral nuclear particles, affecting the mass of atoms and the physical characteristics of different atomic isotopes. HS-C4.4, HS-C6.2, HS-PS1. A1, HS-PS1. A3 Bohr, N. (1913). Atomic and Molecular Constitutions. Philosophical Magazine (London), Series 6 (26), 1-25. Chadwick, J. (1932). Possible Existence of Neutron. Nature, 129(3252), 312. Dalton, John (1805). On The Infusion of Gases by Water and Other Liquids. Memoir of the Manchester Literary and Philosophical Society, Series 2(1), 271-287. Einstein, A. (1905). New Determination of Molecular Dimensions. Annalen der Physik, Series 4(19), 289-306. Einstein, A. (1905). Does the body's inertia depend on its energy content? Annalen der Physik, Series 4(18), 639-641. Einstein, A. (1905). On electrodynamics the body moves. Annalen der Physik, Series 4(17), 891-921. Einstein, A. (1905). At a heuristic point of view on the production and transformation of light. Annalen der Physik, Series 4(17), 132-148. Einstein, A. (1905). On the motion of small particles suspended in the liquid at rest required by the theory of thermal molecules. Annalen der Physik, Series 4(19), 371-381. Planck, M. (1903). Treat thermodynamics. Ogg, A. (trans.). London: Longmans, Green & Co. Adrian Dingle, B.Sc., Anthony Carpi, Ph.D. Atomic Theory II Visionlearning Vol. CHE-1 (3), 2003. Top Page 7 Theory and Atomic Structure by Adrian Dingle, B.Sc., Anthony Carpi, Ph.D. As discussed in our Atomic Theory II module, in late 1913 Niels Bohr facilitated the leathing to new paradigraphs of atomic theory - quantum mechanics. Bohr's new idea that electrons are only available in determined orbits, quantified is revolutionary (Bohr, 1913). As consistent with all the new scientific findings, the new way of thinking about the universe at the atomic level will only lead to more questions, the need for additional experiments and a collection of evidence, and the development of expanded theories. Thus, in the early second decade of the 20th century, another rich vein of scientific work would be mined. In the late 19th century, the periodic table father, Russian chemist Dmitri Mendeleev, had determined that elements could be grouped together in a way that showed gradual change in their observed nature. (This is discussed in more detail in Our Module Periodic Element Table.) By the early 1920s, other periodic trends, such as the number of atoms and ionising energy, had also been well established. Periodic Table of German Physics Elements Pauli made a quantum lege by realizing that there was a difference in ionizing energy and the amount of atoms among atoms with lots of electrons, there should be a way that electrons are not all placed at the lowest energy level. If the atoms of various electrons have all their electrons placed at the lowest energy level, then a very different periodic pattern will result from what is actually observed. However, before we get to Pauli and his work, we need to establish some more fundamental ideas. The development of early quantum theory leans on the concept of duality of wave particles. The simultaneous and complex idea of this is that light (as well as other particles) has properties consistent with both waves and particles. The idea was first seriously oversanned in relation to light in the late 17th century. Two camps are formed over the nature of light: one favors light as a particle and one favors light as a wave. (See Light I: Particles or Waves? module for more information.) While both groups presented effective arguments backed by the data, it was not until a few hundred years later that the debate was resolved. Interactive Animation: The atomic and ionic structure of the first 12 elements Of the late 19th century wave particle debate continues. James Clerk Maxwell, a Scottish physicist, developed a series of similarities that accurately describe light behavior as electromagnographic waves, as if binding on the debate in favor of the waves. However, in the early 20th century, both Max Planck and Albert Einstein conceived experiments showing that the behavior exhibited light that was consistent with it being particles. In fact, they developed theories that suggested that light was a wave particle - a hybrid of both properties. At the time of Bohr's watered-down paper, the time was right for the expansion of this new idea of wave-particle duality in the context of quantum theory, and in French physics crossed Louis de Broglie. In 1924, de Broglie published a PhD thesis (de Broglie, 1924). He proposed an extension of twolity muscle waves light to all things, but particularly to electrons. The starting point for de Broglie is Einstein's equation describing the nature of two photos, and he uses analogy, supported by math, to get an equation known as the de Broglie wavelength (see Figure 1 for visual representation of the wavelength). The equation of the wavelength of de Broglie is, in the great scale of things, a very simple one associated with two variables and a constant: momentum, wavelengths, and constant Planck. There has been support for de Broglie's idea since it made common sense, but the nature of science demands that good ideas and finally demonstrated by experimentation. Experiment. de Broglie doesn't have any experimental data, so the idea remains unconnered for several years. Figure 1: Two wavelength representations of the Broglie (blue line) use hydrogen atoms: radial views (A) and 3D (B) views. It wasn't until 1927 that the hypothesis of de Broglie was shown through davisson-Germer's (Davisson, 1928) experiment. In their experiments, Clinton Davisson and Lester Germer shot electrons on a piece of nickel metal and collected data on observed contradiction patterns (Figure 2). Electron solution patterns are entirely consistent with patterns that have been measured for X-rays and, since X-rays are known to be electromagnetic radiation (that is, waves), the experiment confirms that electrons have wave components. This confirmation means that the hypothesis of de Broglie is correct. Figure 2: Experimental drawings carried out by Davisson and Germer where they shot electrons on a piece of nickel metal and observed a pattern of contradictions. Image © Roshan220195 Interestingly, it was the effort (experimental) of others (Davisson and Germer), leading to de Broglie winning the Nobel Prize in Physics in 1929 for the discovery of the theory of electron wave properties. Without evidence that the Davisson-Germer experiment provided, the 1924 hypothesis de Broglie would remain just that - hypothetical. The sequence of these events is a quintessential example of the theory compiled by experimental data. The Theory of Understanding Inspection must be supported by In 1926, Erwin Schrödinger came from his now-famous equation (Schrödinger, 1926). For about 200 years before Schrödinger's work, the incessant easier F = ma (Newton's second law) was used to describe particle motions in classic mechanics. With the emergence of quantum mechanics, a truly new equation is needed to describe the properties of subatomic particles. Since these particles are no longer thought of as classic particles but as a wave of particles, the partial differential equation of Schrödinger is the answer. In simplest terms, just as Newton's second law describes how the movement of physical objects changes with changing conditions, the Schrödinger equation describes how the wave function (Ψ) quantum system changes over time (Equation 1). Schrödinger's equation was found to be consistent with electron descriptions as waves, and to predict the parameters of the atomic energy levels proposed by Bohr. Equation 1: Schrödinger equation. Schrödinger's equations may be most commonly used to determine the three-dimensional area of space where the electrons given are most likely found. Each area of space is known as atomic orbit and is characterized by a set of three quantum numbers. This number represents the value that explains the Atomic orbits: include its size (n, primary quantum number), shape (l, quantum angle number or azimuthal), and orientation in space (m, quantum magnetic number). There is also the fourth exclusive quantum number for certain electrons rather than specific orbits (s, quantum numbers of spins; see below for more information). Schrödinger's equation allows the calculation of each of these three quantum numbers. This equation is a critical piece in the puzzle of quantum mechanics, as it brings quantum theory into a sharp focus through what amounts to a mathematical demonstration of Bohr's basic quantum idea. Schrödinger's wave equation is important because it bridges the gap between classic Newtonian physics (which is damaged at the atomic level) and quantum mechanics. Schrödinger's equation is considered a monumental contribution to the progress and understanding of quantum theories, but there are three additional, detailed considerations below, which must also be understood. Without this, we will have an incomplete picture of our incomplete understanding of electron relativistics in the atom. German mathematician and doctor Max Born makes a very specific and important contribution to quantum mechanics associated with the Schrödinger equation. Born takes on the wave function produced by Schrödinger, and says that the solution to the equation can be interpreted as a three-dimensional probability of a map where electrons are probably most likely to be found around the atoms (Born, 1926). This map has been known as orbit s, p, d, and f (Fig. 3). Figure 3: Based on birth theory, this is a three-dimensional representation of the electron location around the atom. Four orbits, in increasing complexity, are: s, p, d, and f. Additional information is provided about quantum magnetic orbital number (m). Image © UC Davis/ChemWiki During the year following the publication of Schrödinger's work, German physicist Werner Heisenberg published a paper outlining the principles of his uncertainty (Heisenberg, 1927). He realized that there were limitations to the extent to which electron momentum and its position could be reflected. Heisenberg's principle of Uncertainty places limits on simultaneous accuracy knowing the position and momentum of the particles: As the certainty of one increases, then the uncertainty of the other also increases. The important thing about the principle of uncertainty is that it fits into a quantum mechanical model where electrons are not found in very specific orbits, such as planets - the original Bohr model - and it also dovetails with a map of Born probability. Both donations (Born and Heisenberg's) are taken together with a solution to Schrödinger's equation, revealing that the position in the atoms can only be predicted accurately in a statistical manner. I mean, we know where electrons are most likely found in the atoms, but we can't really be sure of its exact position. Checkpoint The Principle of Heisenberg's uncertainty about the position and momentum of the particles states that as a certainty of one hike, __ another hike. In 1922 German physicist Otto Stern, a Born aide, and Walther Gerlach conducted an experiment in which they passed the silver atom through the magnetic field and observed the sticker patterns. In simple, the results produce two different possibilities associated with a single, 5s valve electron in each atom. This is an unexpected observation, and implies that one electron can take two very different states. At the time, no one could explain the phenomenon that experiments had shown, and it took some scientists, worked independently and in unison with earlier experimental observations, to work within a few years. In the early 1920s, bohr quantum models and various spectras that were produced could be adequately portrayed by the use of only three quantum numbers. However, there are experimental observations that cannot be explained through only three mathematical parameters. In particular, as of 1896, Dutch physicist Pieter Zeeman noted that a single valve electron present in sodium atoms could produce two different spectrum lines with the presence of magnetic fields. This same phenomenon is observed with other atoms with the odd number of valve electrons. These observations are problematic because they fail to adjust the work model. In 1925, Dutch physicist George Uhlenbeck and his graduate student Samuel Goudsmit suggested that this odd observation could be clarified if electrons had corner momentum, the concept that Wolfgang Pauli later called a spin. As a result, the existence of a fourth quantum number has been revealed, which is free of orbit where electrons live, but unique to individual electrons. Considering the rotation, observations by Stern and Gerlach make sense. If electrons can be thought of as a rotating body, the electricity is charged, it will create its own magnetic moment. If the electron has two different orientations (one right hand and one hand left), it will produce two different 'rotations,' and these two different states will explain the anomaly behavior cited by Zeeman. This observation means that there is a need for a fourth quantum number, ultimately known as a quantum number of spins, to fully describe the electrons. Then it is determined that the spin number is indeed required, but for different reasons - whether a fourth quantum number is required. Some Understanding Checkpoints Observations could not be explained mathematically using three parameters because in 1922, Niels Bohr visited his colleague Wolfgang Pauli in Göttingen where he worked. At the time, Bohr was still wrestling with the idea that there was something important about the number of electrons found in 'closed shells' (shells that had been filled). In his own account then (1946), Pauli describes how to build a Bohr idea and draw inspiration from the work of others, he suggested the idea that only two electrons (with the opposite rotation) should be allowed in any one quantum state. He called this 'two values' - the translation of German zweideutigkeit (Pauli, 1925). The consequence is that when a pair of electrons occupies a low quantum of energy (orbital) conditions, any subsequent electrons should enter higher quantum states of energy, also limited to couples at every level. Using this idea, Bohr and Pauli were able to build models of all atomic electronic structures from hydrogen to uranium, and they found that their predicted electronic structures matched known periodic trends existing from periodic tables - the theory met experimental evidence again. Pauli eventually formed what is known as the principle of exclusion (1925), which used the fourth quantum number (introduced by others) to distinguish between the two electrons that make up the maximum number of electrons that could be at any given quantum level. In its simplest form, Pauli's exclusion principle states that no two electrons in the atom can have the same set of four quantum numbers. The first three quantum numbers for any of the two electrons can be the same (which puts it in the same orbit), but the fourth number must be either +1/2 or -1/2, that is, they must have a different 'spin' (Fig. 4). This is what Uhlenbeck and Goudsmit suggest, following the original publication of his theories. Figure 4: Fourth quantum number model, rotation (s). Shown here are models for particles with a rotation (s) 1/2, or half the momentum of the angle. The period described here is rich in the development of quantum theory of the atomic structure. Literally dozens of individuals, some mentioned throughout this module and others do not, contributing to this process by providing theoretical insights or experimental outcomes that help shape our understanding of the atoms. Many individuals worked in the same laboratory, collaborated, or communicated with each other during that period, allowing for rapid evacuation of ideas and refinement that would form modern physics. All these contributions can certainly be seen as an additional building process, in which one idea leads to the next, each adding to the subtle and understanding, and advancing the science field. 20th 20th is a rich period by advancing our knowledge of quantum mechanics, forming modern physics. Researching developments at this time, this module includes ideas and refining built on bohr soil rupture work in quantum theory. Donations by many scientists explain how theoretical insights and experimental decisions revolutionize our understanding of atoms. Concepts include the Schrödinger equation, Born's three-dimensional similarity map, Heisenberg's uncertainty principle, and electron spin. The Main Concept of Electrons, such as light, has proven to be a wave zarah, exhibiting the behavior of both waves and zarah. Schrödinger's equations explain how the wave changes with the time in fashion is similar to the way Newton's second law describes the proposal of classical zarah. Using quantum numbers, one can write wave functions, and look for solutions to equations that help determine the position of the most likely electrons in an atom. Max Born's interpretation of the Schrödinger equation allows the construction of a three-dimensional bar map in which electrons can be found around atoms. These 'maps' have been identified as orbits s, p, d, and f. Heisenberg's Uncertainty Principle stipulates that the position and momentum of electrons cannot be known precisely together, on the contrary we can only think of statistical possibilities of electron location. The discovery of electron spinsrifies the fourth quantum number that is free of electron orbit but unique to electrons. Pauli's exclusion principle states that no two electrons of the same rotation can occupy the same orbit. HS-C1.4, HS-C4.4, HS-PS1. A2, HS-PS2. B3 Bohr, N. (1913). About the constitution of atoms and molecules. Philosophy Magazine (London), Siri 6, 26, 1-25. Born, M. (1926). 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Philosophy magazines London, Edinburgh and Dublin (6th series), 48(286), 719-736 Adrian Dingle, B.Sc., Anthony Carpi, Ph.D. Atomic Theory III Visionlearning CHE-3 (6), 2015. Top Page 8 Theory and Structure of Atoms by Adrian Dingle, B.Sc., Anthony Carpi, Ph.D. In Atomic Theory III: Wave-Partirah Duality and Electron, we discuss progress made by Schrödinger, Born, Pauli, etc. in the use of quantum models to atomic theory. Schrödinger's equation is seen as a major mathematical link between the theory and the use of quantum models. Born takes on the wave function produced by Schrödinger and says that the solution to the equation can determine the energetic and electron positions most likely in the atoms, thus allowing us to build a more detailed description of where electrons may be found in the atoms. This module explores further this solution, electron positioning, atomic orbital form, and the implications of these ideas. As we saw in previous readings, electrons are not the correct particles, but wave particles similar to photos. Since we measure the position of the object with light, the small size of the electrons introduces challenges. If we shine a beam of light on a moving tennis ball, the light has little effect on the tennis ball and we can measure both position and momentum with a high degree of accuracy. However, the electrons are so small that even one photo will influence its trajectory - therefore if we shine a beam of light on it to measure its position, the energy of the photo will affect its momentum, and vice versa. Werner Heisenberg developed the principle to describe this uncertainty, called due to the Heisenberg Uncertainty Principles (Heisenberg, 1927). It tells us that mathematically, the product of uncertainty in position (Δx) and uncertainty in electron momentum (Δp) should not be less than Planck continuously ħ/2 (Equation 1). Equation 1: The Principle equation of Heisenberg's uncertainty, in which Δx is a product of uncertainty in position, Δp is uncertainty in electron momentum, and ħ/2 is a reduced Planck continuous. (Equations are created with CodeCogs online tools.) This is a very small number that can usually be ignored, but when dealing with particles as small as electrons, it is important. Therefore, it becomes necessary to describe the electron position in terms of probability, rather than absolute certainty. We need to say that electrons may be available in atoms in certain areas of high probability, but we cannot 100% definitely position it is accurate. The Checkpoint of Heisenberg's Uncertainty Principles tells us that Because electrons are not true particles, we cannot traditionally describe its movements or locations. In other words, those that will usually be defined by x are simple, y, and coordinates z. Challenges by a combination of wave-wave duality and Heisenberg Heisenberg uncertainty simplified and specified by the Schrödinger equation that, when resolved, produces the function of the waves shown by the Ψ. When Ψ is square, the resulting solution provides the probability of finding electrons in a particular place in the atom. The Ψ2 describes how electron density and the probability of electrons are distributed in space around the atomic nucleus. The application of the Schrödinger equation is easiest to understand in case of hydrogen atoms because the atom of one electron allows us to avoid complex interactions of various electrons. The free time form of the Schrödinger equation (Equation 2) can be resolved to provide a solution that corresponds to the level of energy in the hydrogen atom. In solving this equation, we find that there are various solutions Ψ that we call Ψ1, Ψ2, Ψ3, etc. Equation 2: Schrödinger's equation time is free for hydrogen energy levels. (Equations are created with CodeCogs online tools.) Each of these solutions has different energy that matches what we think of as a different level of energy, or electron shell, in the atom. The lowest distribution of energy electrons is the closest to the nucleus, and it is called soil conditions. Soil conditions are the single most stable electron state in hydrogen atoms. Other higher energy states exist and are called excited states. As each level of energy increases above ground conditions, we refer to them as the second, third, fourth, etc. energy levels, respectively. Checkpoint Energy Understanding electron shell ____ as you move from the atomic nukleus. Acceptable solutions for wave equations for electrons cannot be just any value, but rather they are limited to those who adhere to certain parameters. The parameters are described by a set of three quantum numbers named the main quantum number, the quantum number of azimuthal, and the quantum magnetic number. These quantum numbers are given n symbols, l, and m, respectively. The primary quantum number n is a positive integer starting with a value of 1, in which 1 corresponds to the first shell (lowest energy) known as the soil conditions in hydrogen. The main quantum number then increases with the growing energy of shells (the conditions excited in hydrogen) to the value of n = 2, n = 3, n = 4, etc., as one moves further from the nucrete to higher energy levels. Quantum number azimuthal l (also called quantum number of orbital angle momentum) determines the physical shape, three orbital dimensions in any subshell. The l value depends on the primary quantum number n, and there is a variety of l values for each value n. Value l is a positive integer or 0, and is determined by pushing the integer from corresponding n. For example, when n = 3 we can subtract 3, 2, 2, 1 from n to the revenue value for l from 0, 1, and 2. Therefore, the shell of the soil conditions in the hydrogen (main quantum number 1) has only one subshell, shell 2 has s and p subshells, etc. The subshells indicated by l are also given the designation of letters s, p, d, and f where l = 0 corresponds to subshell, l = 1 p subshell, l = 2 a d subshell, and l = 3 subshell f. Each subshell has a unique shape in the 3D space. Quantum magnetic number m, defining orientation (i.e., position) and number of orbital in any given subshell. The value m depends on the value of l, the quantum number of azimuthal, and m can take the value of +1, -1, and all integers (including 0) in between. So for example, when l = 1, m can be +1, 0, or -1, producing three separate orbitals, on the axe x, y, and z in space. So in the case of l = 1 (p subshell), there are three orbits oriented in different directions in space. Therefore, each energy level is given a unique set of quantum numbers to fully describe it. Only a certain value for quantum numbers in any one level of energy is allowed, and the combination is summarized in Table 1. Analysis of all allowed solutions for wave equations indicates that orbitals can be grouped together in sets according to their l values: s, p, d, and f sets. Table 1: The authorized solution for the wave equation indicates that orbitals can be grouped together in sets according to their l values (i.e., s, p, d, and f set). n l m (principal quantum number) azimuthal quantum number, with letter designations) (magnetic quantum number, with letter designations) 1 0 (s) 0 (s) 2 0 (s), 1 (p) 0 (s), -1/0/+1 (p) 3 0 (s), 1 (p), 2 (d) 0 (s), -1/0/+1 (p), -2/-1/0/+1/+2 (d) 4 0 (s), 1 (p), 2 (d), 3 (f) 0 (s), -1/0/+1 (p), -2/-1/0/+1/+2 (d), -3/-2/-1/0/+1/+2/+3 (f) The first set of sub-orbitals are described by a type of wave function whose probability density depends only on the distance from the nucleus, and whose probability is the same in all directions. Therefore, they have a spherical shape and are called orbits. The function of this wave is all found to have l value = 0 and therefore the value of m = 0, and each energy level has such wave function starting with 1s, and moving to 2s, 3s, etc. (Figure 1). Figure 1: Spherical-shaped spherical orbit. Images © UC Davis ChemWiki The second type of wave function has a density of probability that depends on both distance from the nucleus and orientation together either x-, y-, or z-axis in space. This leads to three, separate, equivalent (or deteriorating) functions of waves that have a diagram shape of eight in the 3D space. This wave function is all found to have l value = 1 and can take three, different m values. This p orbital and exist for every stage of energy except the first, thereh appears as 2p, 3p, 4p, and others (Rajah 2). Rajah 2: rajah eight p orbital. Orbital. © UC Davis ChemWiki The third type of wave function has a propriety that depends on both the distance from the nucleus and orientation along the two x-, y-, and z-axe packs in space. This leads to five different and equivalent wave functions. When wave functions have equal energy, they are labeled slumped. They have a complex shape in a 3D space. This wave function is all found to have a value of l = 2 and can take five different m values. This is called d orbital, and every stage of energy except the first and second has a wave function such as (Rajah 3). Rajah 3: Orbital d, the beginning of a more complex form of orbit. UC Davis © image ChemWiki The fourth type is an orbital set with more complexity in terms of wave function, position, and shape in 3D space. This wave function is all found to have a value of l = 3 and can take seven different m values. They, too, are all slumped. This is called f orbital and every stage of energy except the first, second, and third has such wave function, starting with 4f and moving to 5f (Rajah 4). Rajah 4: Orbital f, which continues the complexity of the form as seen in orbit d. image © UC Davis ChemWiki To formulate, the orbits contained in the first four stages of power are as follows in Schedule 2: Schedule 2: Orbital relating to the first four stages of power. n orbit 1 1s 2 2s, 2p 3 3s, 3p, 3d 4 4s, 4p, 4d, 4f Main Quantum Checkpoint Understanding number n Because orbit is determined by a very specific solution to the Schrödinger equation, electrons must absorb a very specific quarry of energy to be promoted from one stage of energy to another. By exposing electrons to external energy sources, such as light, it is possible to promote electrons from the state of their soil to other potentially higher energy-powered, so-called spoken countries. These 'teruja' electrons quickly return to a lower power position (to regain stability associated with lower energy), and in doing so they release energy in a certain frequency corresponding to the power difference between the electron's orbit, or the grip. Light power is related to frequency (f) and planck pemalar (h) in the equation E = hf. Electrons in hydrogen atoms will absorb a certain frequency of energy corresponding to the energy gap (ΔE) between two levels of energy, while allowing the promotion of electrons from a rather low energy level to a rather high energy level. Examining the absorption spectrum produced when hydrogen is disindir, we note the line pattern that supports this (Rajah 5). Rajah 5: Release of hydrogen and spectra absorption from the series Image © Chem1 Virtual Text Tool, which is sapled from Online Travel via the Bohr Astronomy site suggesting that electrons can turn to but only in scripted, the measures are defined. He argues that electrons are restricted to certain areas of space around the nucleus, just as planets in our solar system are confined to certain routes. The spectrum of discrete line hydrogen absorption (rather than a continuous spectrum) suggests that only a very specific transition can be made and is a testament to quantum models. In addition to the absorption spectrum indicated above, other spectrums of the same can be observed, this time with the energy produced rather than absorbed. In this case, the electron falls back from a higher calm to a lower dry than promoted as in the absorption spectrum. This is called the atomic relief spectrum and again consists of discrete, individual lines. Checkpoint Understanding When energy is removed, the electrons Each series of lines appearing in the spectra are named after their findings. The Lyman series is the lines where electrons are either promoted from, or to, orbit 1s. In the Balmer series, Paschen, Brackett, and Pfund, electrons move either from or to orbit where n = 2, 3, 4, and 5, respectively. Different energy gaps (and therefore different frequencies) correspond to the lines in different areas of the electromagnetic spectrum as a whole (Fig. 6). Lyman lines are in the ultraviolet region, balmer in the visible, and Paschen, Brackett, and Pfund in various parts (near and far) inframerah. Figure 6: Diagram of energy level transition for hydrogen. In hydrogen atoms, and other single electron species such as He+ and Li2+, it is found that orbital energy depends only on their distance from the nuclus, with increased energy as n increases. The lowest energy orbits are 1s, followed by 2s and 2p (which have the same energy), and 3s, 3p, and 3d (which have the same energy) etc. Therefore, the soil conditions (lowest energy states) for hydrogen atoms with only one electron are with electrons living in orbit 1s. The energy difference (ΔE) required to promote electrons from one level to another can be determined by using one version of the Rydberg Formula, where RH is constant for hydrogen and ni and nf is the early and last energy levels of each. Equation 3: Rydberg Formula (created with CodeCogs online tool.) Since other single electron species such as He+ and Li2+ contain more protons than hydrogen atoms, electrons experience greater attraction from the nuclus. Therefore, different, a larger amount of energy is required for the promotion of electrons in this species and a modified Rydberg marketer is required in the calculation of ΔE for a single electron species other than a hydrogen atom. CheckPoint In atoms that have only one electron, the orbital power dependent on hydrogen atoms is the easiest to learn because a single electron is both not influenced by other electrons and does not affect any other electrons. In this case, the wave function is rather easy to guess. However, in more complex atoms the calculation is not so easy. For example, when two electrons are present, such as in helium, things become much more complex. Rather than just one potentially easy power pull between the nucleus and a single electron in a hydrogen atom, in helium there are now three potential energy terms to consider: the attraction between the nucleus and each electron, and the penetrater between the two electrons. The addition of more electrons leads to equations and calculations becoming increasingly complex. The complexity of Schrödinger's equations under many electron circumstances means it needs to be solved by a series of budgets rather than directly. One method is to effectively use a single electron system over and over again to generate what amounts to a budget answer. Although not entirely correct, such a budget proved to be highly functional and produced a reasonable answer to the problem of various electrons. The budget works well to once again produce A2 which predicts the propriety of electrons in the atomic space of various electrons. Orbits in the state of single electrons and those in various electrons are still determined by the same set of quantum numbers, but a difference arises regarding the specific calmness of the orbit. Previously, energy depended only on the distance from the nucleus (i.e. n energy in the system of various electrons was found to depend on the value of l). This complicates the idea of associating electron power solely to distance itself from the nucleus (i.e., orbital power following the pattern of 1s < (2s = 2p) < (3s = 3p = 3d), etc.). Instead, it produces a relatively calmed following sequence for orbit. 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p, and others. Note the insertion of 4s and 5s earlier in sequence than expected. This order is indicated on the periodic schedule. When we reach the end of the three-term period on the periodic table (element 18, argon), we will charge 1s, 2s, 2p, 3s, and a 3p orbit giving the number (as expected) of 18 electrons. One might expect the next orbit to be filled into 3d, because after all, they are also at the third energy level. However, the next elements on the table periodically, potassium and calcium (elements 19 and 20, respectively), fill their 4s orbits before their 3d. Moving further along the fourth period for element 21, the scandium, we find that subshell 3d is now beginning to populate. If the subshell is full on element 30, zink, we then fill the subshell starting with element 31, gallery. In summary, regularly booked schedules in such a way that it reflects the relatively calming sequence of orbits given above. As seen before, the concept of electron rotation introduces the need for the fourth and final quantum number, the quantum number of the rotation, s. Its introduction ensures that each electron has a unique set of quantum numbers according to the Pauli Exclusion Principle. With only two values for s (+1/2 or -1/2), it follows that each orbit can only hold a maximum of two electrons. The notation used to describe the electronic configuration of atoms involves the use of superscripts to indicate the number of electrons in any given orbit. For example, in a beryllium case that has a number of four electrons, 1s2 2s2 indicates that both 1s and orbital 2s are filled with two electrons each. When we identify the orbit in which electrons in an atom occupy, we generally follow the principle referred to as the Aufbau principle, where the lower energy orbit is filled first. So, for example, 1s is filled before 2s, and 2s before 2p, and others. But what about degenerate orbits, those with the same energy, as the three orbits that exist at sub-levels of 2p? Hund rules state that when electrons are placed into any set of degenerate orbits, one electron is placed into each orbit before any rotational pair takes place. This means, for example, if we assume three electrons are put into a 2p degenerate orbit, one would look at the electronic configuration of 2px1 2py1 2pz1 (remember orbit-oriented x, y, and z axes, respectively) before we see any pair. The fourth electron to enter the 2p sub-shell will realize the need for an inevitable pair, therefore 2px2 2py1 2pz1 (Rajah 7). Rajah 7: Hund Regulation Illustration shows the placement of electrons in various orbits of nitrogen (N) and oxygen (O). Now that we have a sense of electron configuration, their power, and their most likely position in atoms, the next step is to describe the behavior of electrons when atoms interact with each other. The interaction of electrons forms the basis of a major chemical process called chemical bonds. And understanding the behavior of electrons gives us a better understanding of the chemical behaviors of the elements and how they are. Our Atomic Theory series continues, exploring quantum models of atoms in more detail. This module looks closely at schrödinger equations thatrify the power and possible position of electrons in atoms. Using hydrogen atoms as an example, the module explains how orbit can be explained by the type of wave function. Evidence for orbit and quantum models is provided by spectra absorption and release of hydrogen. Other concepts electrons, the Aufbau Principle, and the Hund Government. The Main Concept of electron wave zarah properties means meaningful Their position and momentum cannot be explained in easy physical terms but must be explained by the wave function. Schrödinger's equations explain how the wave changes with the time in fashion is similar to the way Newton's second law describes the proposal of classical zarah. This equation allows the calculation of every three quantum numbers related to the orbit of individual atoms (principal, azimuthal, and magnet). Heisenberg's uncertainty principle stipulates that the position and momentum of electrons cannot be properly known together; otherwise we can only guess at the statistical probability of electron location. The discovery of electron spinsrifies the fourth quantum number that is free of electron orbit but unique to electrons. Pauli's exclusion principle states that no two electrons of the same rotation can occupy the same orbit. Quantum numbers, when taken as a set of four (principal, azimuthal, magnet and rotation) describe an acceptable solution to the Schrödinger equation, and thus, describe the position of electrons most likely in atoms. Orbitals can be thought of as three-dimensional areas of space, determined by quantum numbers, which describe the most likely electron position and energy in atoms. HS-C1.4, HS-C4.4, HS-PS1. A2, HS-PS2. 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Top Page 9 It's almost the beginning of the school year, and you've gathered 10 friends for the bonfire end of the semester. What would bonfire without more? You wrap up the supply, making sure you have enough to make another one for everyone. You can think of making more as chemical equations (see our Chemical Equations module for more on them): 2 Graham keropok + 1 piece of chocolate + 4 mini-marshmallows → 1's more Just like with chemical equations, multipliers in the presence of pegs and products show a peadarant in which they respond to more so. So, to make 10's more, you need to: 20 Graham crackers + 10 pieces of chocolate + 40 mini-marshmallow → 10's over Congratulations, you've just made it through your first workout in what chemists call stoichiometry. The mouth of the term was coordinated in the 1790s by Jeremias chemist Benjain Richter, who was fascinated by probable mathematics combining chemicals, convinced that it held clues to the nature of things (which it indeed was; Dalton drew this math to devise his initial atomic theory, described deeper in Our module Of Early Ideas Matter: From Democritus to Dalton Preliminary Ideas on Things: From Democritus to Dalton Richter combining the stoicheion of the Greek word, meaning elements, and metrons, which mean measurements. In other words, stoichiometry is a way to measure the amount per reactan that combines chemical reactions, in our case, the number of reactions (20 graham crackers, 10 pieces of chocolate and 30 mini-marshmallows) which in turn predicts the number of products (10 s'mores), or vice versa. Stoichiometry may seem like a complicated word, but it is a relatively simple concept when you apply it to a chemical equation: the proportions expressed in the chemical (cocaine) equation can be used to predict how much the product will be produced from the measurement of the given reaction. For example, we use stoichiometry to determine how much more reactants we need to make another 10. We can also use stoichiometry to predict how much products we will get with the amount of each reacment we have. If we have a lot and lots of chocolate and marshmallows but only 12 graham crackers, how much more can we make? Again, our similarities are: 2 Graham Crackers + 1 chocolate + 4 mini-marshmallows → 1 over If we have 12 graham crackers, that's enough to make 6's more. No matter how much extra chocolate we have, because without graham crackers, it's not over yet. So the ratio of graham cracker moles to more produced is: Using the same mole ratio concept as described above, stoichiometry is used to find out how much reakant is needed to make the desired quantity of products in the laboratory or manufacturing facilities. An important example of industrial is the production of nitrogen-based fertilizers, which provide essential nutrients to the soil and allow modern farmers to expand more food per acre. For centuries, farmers have understood the importance of adding nutrients to the soil where they grow crops, but before the 1900s they were limited to using animal fertilizer or expensive, naturally occurring mineral deposits as fertilizers. In the 1840s, chemists Justus von Liebig identified nitrogen as the main ingredient of steel. However, despite the large amount of nitrogen in the atmosphere, there is no easy way to exchange to the form that can be taken by plants. This all changed in the early 1900s when German chemist Fritz Haber created a chemical process to convert nitrogen to ammonia (NH3), a compound that often gives home cleaners their characteristics, and plants that can be used as a source of nitrogen. His initial method was only economical on a small scale, so Haber worked with a German colleague, Carl Bosch, to adjust this process to work on an industrial scale. The Haber-Bosch process was sometimes referred to as one of the most significant inventions of the 20th century, and it led to Haber winning the Nobel Prize in Chemistry in 1918. In its equational form, the Haber-Bosch process is quite simple: N2 + 3H2 → 2NH3 The ability to perform this simple reaction on a large scale has significant historical consequences. Cheap ammonia provides space for cheap fertilizers that are widely available, which created a boom in agriculture (and a related population increase) in the 20th century. And it is indirectly prolonged World War I by providing Germany with the cheap nitrogen resources needed to make gunpowder. Some scientists have recently questioned whether the Haber-Bosch process is a sustainable practice, given the growing environmental impact of agriculture and population, as well as the fact that great energy is needed to generate hydrogen gas. Spray paddy fields with fertilizer. The manufacture of nitrogen-based fertilizer depends on the stoichiometry to calculate how much the initial material (N2 and H2) is required to produce the desired amount of ammonia (NH3) for use in fertilizer. Image © Jan Amiss Let's request for our stoichiometry discussion here and implies that an agricultural company needs to produce 1,500 kilograms of NH3 to meet fertilizer demand. How many N2 and H2 should they start? Again, the equation is: N2 + 3H2 → 2NH3 Looking at the equation, we see that the N2 mole ratio needed to produce NH3 is: (1 mole N2) / (2 mole NH3) Now we will use this mole ratio to determine how much remensement we need to start making 1,50 First, a reminder: when we calculate the amount of material in the reaction, we need to convert the mass of each material into a mole. Why? Because the ingredients involved do not have the same weight. Think in more: 1 piece of chocolate weighs more than 1 mini-marshmallow. If we use mass in equations rather than the number of pieces, we might say that one needs more 1 gram of chocolate and 4 grams of mini-marshmallows. But in reality, that would amount to one piece of chocolate and about 50 mini-marshmallows! (To about converting from gram to mole, see our module about moles and atomic mass.) So, we know that we want to make 1500 kg of NH3, Let's start by Kilograms to grams as follows 1,500 kg NH3 x 1000 g/kg = 1,500,000 g of NH3 Then, we need to calculate how many moles there are. To do so, we breed the mass of NH3 molecules (17 g per mole) with the number of grams, setting the equation up to gram cancel and the answer is in moles. We see that: 1,500,000 g NH3 × (1 mole NH3) / (17 g NH3) = 88,325 moles NH3 Next, we can use mole ratio to find out how many N2 moles are needed. Since we need 1 N2 mole to produce 2 NH3 moles, we use that mole ratio to determine how many N2 moles are needed to produce 88,235 NH3 moles: 88,235 moles NH3 × (1 mole N2) = 4,4 N44,117, N44,117, N44,117, N44,117, N44,117 NH3 moles) = 44,117 moles using the mass of N2 molecules to find out how many grams N2 is needed, then convert N2 grams to kg N2, because that's the unit we want for our answer: 44,117 N2 moles × (2 8 g N2) / (1 mole N2) = 882,340 g N2, or882.340 kgNow that we know how many kilograms of N2 we need, we can use the masse mole ratio (N2 and H2) to find out how many H2 moles are needed. Remember, the equation states: N2 + 3H2 → 2NH3 Mole ratio for H2 to N2 is 3 to 1. So, for each nitrogen mole, we will need three times more hydrogen moles: (3 H2 moles) / (1 mole N2) Remember, we need to calculate how many hydrogen moles are needed, and then convert the mole into a gram of hydrogen. Using nitrogen moles we calculated above, we can get kg H2: 44,117 moles N2 x (3 H2 moles) / (1 mole N2) × (2 g H2) / (1 H2 mole) = 264,702 g H2, or 264.702 kg This is an important information for the manufacturer Although the existing nitrogen is from the air, hydrogen gas is not. So manufacturers may need to buy hydrogen gas, which is expensive to generate, potentially explosive, and difficult to transport and store. Therefore, manufacturers need to know exactly how much hydrogen gas is needed. In the above case, the manufacturer will have an unlimited amount of nitrogen gas, but the exact amount of hydrogen gas. Therefore, the amount of hydrogen gas will limit the amount of ammonia that can be made (similar to the number of graham crackers can limit the number of more that can be made). We will say that hydrogen is a limited reactan, which means that this is a reaction to be used first. As a result, the amount will determine how many products are produced. Determining how much reactory is needed to produce a certain amount of products is one of the most important stoichiometry applications. We will describe this first with more Tell me you have the following amount of more reaktan: 120 graham keropok 70 pieces chocolate 200 mini-marshmallows Here, again, is more equation: 2 Graham keropok + 1 chocolate chip + 4 mini-marshmallows → 1's more How How To s'mores can you make from your reaction? That will depend on a limited reaction, which will be exhausted first. To determine with reacment is limiting, you need to first calculate how much more you can make with each reacment. You can do this using a mole ratio: Limiting reaction is an important concept in any manufacturing process. Manufacturers know they want to make a certain number of products, and will buy reacments accordingly. In many cases, it is more economical to make the most expensive reactions be a limited one, reducing the cost of surup and waste. Silver nitrite is a good example. These compounds, AgNO3, have been used since ancient times as disinfectant agents and wound healing. Today it is used in bandages and other medical applications, as well as water purification. It can be easily made by reacting pure silver with nitric acid, according to the equation: 3Ag + 4HNO3 → 3AgNO3 + 2H2O + NO Silver is a more expensive reaction of nitric acid, so someone who uses it to produce silver nitrate may want to make limited silver reactan. By starting with the specified amount per crack, you can specify not only a limited reaction but also the mass of the product to be produced and the amount of reacment that remains excessive. Let's say we start with 150g of silver and 150 g of nitric acid. How much AgNO3 can we make, and whose reaction is limited? To find answers takes a few steps: 1) convert each reactan to 2 moles) using a mole ratio to determine how many reactive one moles are required to use 3) others calculate the number of products based on using all the limited reactions. Step 1: Switch to a mole 150g Ag × (1 mole) / (108 g Ag) = 1.39 mole Ag and 150g HNO3 × (1 mole) / (63 g HNO3) = 2.38 mole HNO3 Step 2: Use a mole ratio to eanning Ag moles to HNO3 moles (4 HNO3 moles) / (3 moles Ag) Because silver is our expensive reaction, we want to use it all. We can calculate how many HNO3 moles are needed to respond with an entire 1.39 mole Ag, establish an equation so that the HNO3 mole cancels: 1.39 mole Ag × (4 HNO3 moles) / (3 moles ag) = 1.85 mole HNO3 is required To use all, Ag, we need 1.85 HNO3 moles. Take a look at our calculations above. How much HNO3 do we have? We have 2.38 moles - more than we need. In other words, if we put all the two reactions together, silver will be used first, and there will be HNO3 left. That makes silver a limited reacment. This example shows the importance of switching to a mole first. We started the same mass per reactor, 150 g. But the mass does not tell us how many particles there are. That's what the mole unit tells us. Knowing that silver is a limited reactan, we can go further and determine how many moles AgNO3 is produced from 1.39 moles we start with. This time we use the ratio of lahiduk ratio between Ag and AgNO3. In reaction: 3Ag + 4HNO 3 → 3AgNO3 + 2H2O + NO There are 3 AgNO3 moles produced for every 3 Ag moles used. So: 1.39 mole Ag × (3 mole AgNO3) / (3 mole Ag) = 1.379 mole AgNO3 produces such calculations it is important for our ability to efficiently manufacture and use chemicals, as well as our ability to understand the effects of the reactions that occur in the world every day. For example, an engineer for a paint manufacturer must consider the ratio of different chemical moles in paint, which will determine the cost of producing the paint. On a larger scale, stoichiometry plays a role in understanding climate change: if we know the quantity of different types of fossil fuels burned in a year, we can determine how much CO2 has been added to the atmosphere. From planning to streamlining manufacturing and generating environmental data, we can use stoichiometry to predict and plan the results of many chemical processes. Stoichiometry is a chemical mathematics. Starting with a balanced chemical equation, we use the proportional properties of the chemical reaction to calculate the amount of reacment required at the beginning or predict the number of products to be produced. Although it may not seem all that chemical, stoichiometry is an underlying concept of our ability to understand the effects and implications of many chemical processes. Bandage manufacturers can use mole ratio to determine how much silver is needed (and there are costs) to treat a bunch of bandages with silver nitrate. A fertilizer company may use the concept of limiting the reaction to knowing how much products they can produce with the amount of hydrogen gas given. And so on. Stoichiometry, the ratio of fecal ratio, and limiting reacment are much-needed concepts to fully understand any chemical process. Stoichiometry uses the proportional properties of the chemical equation to determine the amount of reactiveness required to produce a certain number of products or predict the amount to be generated from the amount of reactors given. The ratio of lahiduk ratio shows the proportion of one reactive or product in response to another, and comes from a balanced chemical equation. Although we may need to adjust the amount of reaction to produce more products, the ratio of reaction to the product is always equal to a balanced reaction. Limiting reacments are chemicals used in advance in response. it can be determined by comparing the number of reactive moles in the hands and the ratio of moles between the rig and the product in a balanced reaction. Stoichiometry is a chemical mathematics. Get started with balanced chemistry, we use chemical response-based properties to mistaking the amount of reakant required at the beginning or predicting the amount of to be produced. Although it may not seem all that chemical, stoichiometry is an underlying concept of our ability to understand the effects and implications of many chemical processes. Bandage manufacturers can use mole ratio to determine how much silver is needed (and there are costs) to treat a bunch of bandages with silver nitrate. A fertilizer company may use the concept of limiting the reaction to knowing how much products they can produce with the amount of hydrogen gas given. And so on. Stoichiometry, the ratio of fecal ratio, and limiting reacment are much-needed concepts to fully understand any chemical process. Haber-Bosch Reaction: Early Chemical Impact on Sustainability Overview of Haber-Bosch Process: Robin Marks, M.A., Anthony Carpi, Ph.D. Stoichiometry Visionlearning Vol. CHE-4 (8), 2019. Page 10 Physical and Real Estate State by Robin Marks, M.A., Anthony Carpi, Ph.D. This is the latest version of the Water module. For an earlier version, go here. Before we get started, get yourself a glass of water. By the time you reach the end, you will have a greater appreciation for the miracle of this liquid. Got your glass? Now take a sip and think about all the water roles that play in your life. For one thing, your body can't work more than a few days without it. You use water to wash yourself, your clothes, and your car. Water puts fire, cooking our food, making our soaps become sudsy, and hundreds of other things. Water is certainly important for our lives on Earth. Water is so central to our existence that you may be surprised to know that it is a rare and unusual substance in the universe. The water at once is very important and so difficult that ecologists (scientists looking for life outside of Earth) set their sights on the planet where water may exist. Life, it seems, can be difficult in acid, lye, extreme salt, extreme heat, and other conditions that will kill us humans. But it cannot exist without water. Despite its shortcomings across the universe, so much water on Earth is that we are not always aware of how special it is. For starters, water is the only substance that exists naturally on our planet as solid (ice and snow), liquids (rivers, lakes, and oceans), and gas (water in the atmosphere as humidity). As you may remember (or can read about in our module on the States of Things), water molecules are in different energy conditions in each phase. The amount of energy needed to go from solid to liquid and liquid to is related to how water molecules interact with each other. These interactions are also related to how atoms in water molecules interact with each other. Our Chemical Bonding: The nature of the Bon Chemical module discusses how the polished shape stretches across molecules; in the bond between oxygen and hydrogen, electrons are shared uncontectly, drawn slightly more to oxygen. As a result, partial negative charges (ð-) forms at the end of molecular oxygen, and partial positive charges (ð+) forms on each hydrogen atom end (Fig. 1). Figure 1: Dipoles arises in water molecules due to the unearthical sharing of electrons. Since hydrogen and oxygen atoms in molecules carry opposite (albeit partially) charges, nearby water molecules are attracted to each other like small magnets. Electrostatic attraction between hydrogen ð+ (ð means partial charge, value less than electron charges) and ð-oxygen in adjacent molecules called hydrogen tie-ups (Fig. 2). Figure 2: Hydrogen bonds between water molecules. The slightest negative charge over oxygen atoms is attracted to a slightly positive charge on the hydrogen atom. Hydrogen bonds make water molecules stick together. These bonds are relatively weak compared to other types of bloody bonds or ionics. In fact, they are often referred to as attractive force compared to true bonding. Yet they have a huge impact on how water behaves. There are many other compounds that make up hydrogen bonds, but the one between the water molecules is very strong. Figure 2 shows why. If you look at the central molecules in this figure, you see that the tip of the molecular oxygen forms a hydrogen bond with two other water molecules; in addition, each hydrogen on the central molecule is attracted to a separate water molecule. As illustrations show, each water molecule forms an attraction with four other water molecules, a connection network that makes hydrogen bonding in water especially strong and lends material that many of its properties are unique. Now it's time to use that glass of water. If you have some ice cubes, fall one on your glass. You'll notice that it's floating. Its ability to bob to the top of the waterline means that ice (water in its solid state) is less compact than liquid water. (To review density and flotation, see our Density module) This is not a normal situation; If you put most of the solid candle into a melting candle vat, it will sink towards the bottom (and may melt before it gets there). To understand what causes floating ice but solid candles to sink, let's think in advance of what happens when the liquid turns to solid (again, the countries of the Articles module can be useful reviews here). In liquid, molecules have enough kinetic energy to keep moving. When the molecules came close to each other, they were pulled together by the intermolecular team. At the same time, molecules have enough kinetic energy to break free of those forces and be drawn to others nearby. Therefore the flow of liquids due to the intermolecular attraction can be broken down and renewed. Liquid freezes when reduced kinetic energy (i.e. reduced temperature) is enough that the attractive force between molecules can no longer be broken, and the molecules become locked in static lovers. For almost all compounds, lower energy and lack of movement between molecules means that the molecules in solids are packed together more stringently than liquid conditions. This is the case with candles and candles that are so solidly solid out of liquid and sink. In the case of water, though, the shape of the molecules and the strength of hydrogen bonds affects the order of molecules. In liquid water, the hydrogen bond draws close molecules together. When the water freezes, the dipole ends up with charges such as repeling each other, forcing molecules into a fixed flict where they are further away from each other than they are in liquid water (Fig. 3). More space between molecules makes the ice less compact than liquid water, and hedes it floats. Figure 3: When the water freezes, the same tip is charged dipoles repel each other, pushing the molecules apart. This means that there is more space between molecules in solids than in liquid, making solid (aka, ice) less compact. Water is sometimes referred to as a universal solven, since it dissolves more compounds than other known liquids. The pole of water molecules allows it to easily dissolve other polyuns molecules, as well as ions. (See our Solutions, feasibility and collaborative properties module for deeper discussion of liquidation.) The ability to dissolve the material is one of the properties that make water essential for life. Most biological molecules, such as DNA, protein, and vitamins are poles, and important ions such as sodium and polysysum are also charged. In order for any of these compounds to carry out functions in the body, they need to be able to circulate in the blood and the fluids in and between the cells, all of which are mostly water. Because of its pole, water can dissolve these ingredients and other ingredients, allowing their free movement around the body. Some biomolecules, such as fats and cholesterol, are not polar, and do not dissolve, and do not dissolve in water – however, the body has developed a unique way to distribute and store these substances. Water can also dissolve gases such as oxygen, allowing fish, plants, and other aquatic life to access these dissolved oxygen (Fig. 4). O2 is not a pole molecule; it dissolves because polar charges in water molecules induce disposal in oxygen, making it dissolved and so available for aquatic life. (Learn more about the dipole-dipole interactions that are instilled in our Liquid Real Estate module.) Figure 4: When water and oxygen molecules meet (left), the negative disposal of water repels electrons around the oxygen molecules, creating a temporary dipole in molecules (right). Let's get back to your water glass. Fill the glass only for and stop. Then, slowly, add a little more. You will see that you can actually fill the glass slightly removing its rim, and the edge of the water will be round against the glass, holding the water in. Again, the hydrogen bonds behind this act, resulting in a fusion. Fusion prevails when molecules of the same type are attracted to each other. In water case, molecules form a strong hydrogen bond, which holds the material together. As a result, the water is very high tide, in fact, it is the most tidal of all the thaw instead of metal. The blend applies throughout your glass of water, but it is very strong on the surface. Molecules are less neighbourly (because they don't have on the very surface), and so create stronger bonds with the molecules that are near them. The result is called surface tension, or the ability of the material to resist interference to its surface. Dip your finger into your water glass and then pull it out. The drop that forms at the end of your finger is held together by surface tension. Surface tensions were misunderstood in a voracious debate between Galileo Galilei and his rival chief, Ludovico delle Colombe in 1611. Delle Colombe, a philosophist, is at odds with some of Galileo's ideas, including his explanation that ice floats on water because it is less dense. So the philosophist challenged Galileo to the debate, which Colombe's investigators believe will prove his own intellectual superiority. Delle Colombe championed the idea (incorrectly) that ice floats not because of density, but because of its shape, which it sees as broad and flat, like ice in a tasik. To prove the truth of his theory, he used black wood, which is slightly denser than water, in demonstrations before the audience was curious. He dropped a wooden spherical into the water, and he drowned. He then put a thin wafer of wooden flats on the surface of the water, and he floated. Delle Colombe calls himself a winner. Galileo was disappointed. His observations around the world suggest that his explanation, not Colombe's delle, is correct, but that he cannot explain the results of Delle Colombe's experiments. If he had known about molecules and polished and hydrogen bonds at the time, Galileo would have offered this explanation: When delle Colombe floated the thin black wood cakera, he took advantage of the combined nature of the water and the surface tension arising from it (Rajah 5). When a black wood wafer appears to float on the water, the power charged by its jisim has been circulated to the entire surface of the water below. In other words, a single pinpoint-sized area of surface water only needs to support a pinpoint-sized piece of black wood on it. Hydrogen bonds between water molecules are quite strong supports the weight of the disc. When colombe delle puts the sphere in the water, water, pinpoint-sized areas that first touch the heavy bore water of the entire sphere, which more than the surface strain of the water can support. Galileo's limit was known to this at the time, he could easily disprove Colombe's delle - he simply pushed the wafer through the surface to break the surface strain, wafer would sink. Figure 5: Water molecules on the surface form stronger hydrogen bonds between them than doing molecules across the water. These stronger bonds are responsible for surface tensions. Image © USGS Strain the same surface is what allows the leaves to stay on the surface of the lake and dewdrops to comply with the spider's web. Even some animals take advantage of this phenomenon - Basilisk lizards (Fig. 6), striders water, and some small animals and other bugs appear to run on water by taking advantage of the surface strain of the water. Figure 6: Basilisk lizard (Basiliscus basiliscus) runs on the surface of the water. S1 movies from Minetti A, Ivanenko Y, Cappellini G, Dominici N, Lacquaniti F. Man Walking In Place On Water in Simulation Reduced Gravity. PLOS ONE. DOI:10.1371/journal.pone.0037300 For your next observation, take another sip of water, and pay attention to the glass side. Chances are you'll see some drops that get stuck to it. Gravity attracts these drops, so something else must keep them stuck there. That something else is an adhesive, a water attraction to another type of molecule; in this case, the molecules that form the glass. Because of the molecular pole, the water exhibits stronger adhesives to the surface that has some clean electrical charges, and the glass is one of those surfaces. But put a droppe of water on a non-pole surface, like a piece of candlestick paper and you'll see it taking shape differently than the one it complies with. On candlestick paper, water drops take the right form of drops as there is little adhesive and cohesive force pulling the decrease into the sphere. But on the glass you will see drops flattened and deformed slightly as the force of the adhesive pulls it more to the surface of the glass. Both solidarity and adhesive (Figure 7) occur with many compounds other than water. Pressure-sensitive tapes, for example, stick to the surface as they are coated with high viscosity fluids that adhere to the surface they are pressed on. Generally, you can overcome the force of these stickers interestingly, for example – you can easily lift the Post-it® a note from the page. But sometimes the adhesive power is stronger than the forces holding the surface together - pull the tape from a piece of paper and you remove the pieces of paper with tape. Let's get back to our glass of water, and see in it where the surface of the water fills the glass. Water surface edge Up a little bit in the glass. It is also lekatan – water is provided adhesives with glass. If you have a clear plastic straw, you can put one end into the water and see that the liquid climbs hay a little, on top of the surface of a glass of water remaining. It actually moves upwards against gravity! What happens in your straw is a phenomenon called capilary action (Fig. 7). Capilary action occurs in small tubes, where the surface area of the water is small, and the force of the adhesive—the attraction of water to the pole glass or other material—overcomes the force of unity between the molecules of the surface. Figure 7: The water molecular attraction to the side of the vessel narrow (adhesion, red arrow) is stronger than unity (orange arrows) pulling the water molecule together. The result is a capillary action, in which the force of the adhesive pulls the liquid up (the purple arrow). Another way to look at the effects of adhesives and solidarity is to compare pole fluid behavior and nonpolar. When you put the water in a test tube, the adhesive makes water along the edges moving slightly upwards and creating a concave meniscus. The merzor of the liquid, on the other hand, is not polar and therefore is not attracted to the glass. In test tubes, the solidarity on the mercry surface is much stronger than adhesives to the glass. The surface tension in the mercry forms a convex meniscus, just as the way the water forms a little bulge over your very full glass (Fig. 8). Figure 8: Water and mercry behave differently in test tubes made of polar glass. The water complies with the glass, brings the side up and forms the surface of the concave. Nonpolar merqueries are not attracted to glass. The unity between the merzy atoms creates a surface strain that forms the surface of the convex. Images © USGS adhesives and capilary action are among the playing forces that help plants consume water (and dissolve nutrients) in their roots. Capilary action also keeps your eyes away from drying out, as saline water flows from small channels in the outer corner of your eyes. With each overnight, you spread the water from the channel, and capillary action brings more fluid to the surface. If you want to see capilary action at work, put a few drops of red food dye in your water glass, and then release the stick or two leafy celery into it. After a day or two, your green celery will be immersed with red. Water is a truly unusual and important substance. The unique chemical properties of water that provoke surface tension, capillary action, and low density of ice play an important role in life as we know it. Floating ice protects aquatic organisms and keeps them from being frozen in winter. Capilary action keeps plants alive. Surface tension allows the lily pad to remain on the surface of the lake. In fact, water chemistry complex and important that today's scientists are still working to understand all these shows can perform. Key Concepts Water has several unique properties that make it important in both the chemical and biological worlds. Water molecular poles allow liquid water to act as a universal solvent, able to dissolve many ionic compounds and covalent poles. Water molecular poles also result in strong hydrogen bonds that provoke phenomena such as surface tension, adhesives, and solidarity. Everts, S. (2013). Galileo on ice: Researchers commemorate scientists' debate about why ice is floating on the water. Chemical and Engineering News, 91(34), p.28-29. Heilbron, J.L. (2012). Galileo. Oxford University Press. Lo Nostro, P. and Ninham, B.W. (2014). Aqua incognita: Why ice floats on water and Galileo 400 years old. Connor Court Press. MachLachlan, J. (1999). Galileo Galilei: First physicist. Oxford University Press. Whitehouse, D. (2009). Renaissance genius: Galileo Galilei and his legacy to modern science. Sterling Newspaper. Robin Marks, M.A., Anthony Carpi, Ph.D. Air Visionlearning Vol. CHE-4 (6), 2018. Top Page 11 Physical Conditions and Properties by Anthony Carpi, Ph.D. As a youngster, I remember hanging wondering on a boiling water pot. Seeking an explanation for the ensuing bubble, I believe for the time that the movement of hot water pulls the air into a saucepan, which then buckles back to the surface. Little did I know that what's going on is more magical than I imagine – the bubble isn't broadcast, but actually the water is in the form of gas. The different states of the matter have long confused people. The ancient Greeks were the first to identify three classes (what we call states) things based on their observations of water. But those same Greeks, in particular the Philosopher Thales (624 - 545 BCE), wrongly suggest that because water can exist as solid, liquid, or even gas under natural conditions, it must be a single major element in the universe from which all other substances are made. We now know that water is not a fundamental substance of the universe; in fact, it is not an element. To understand the different states where things can exist, we need to understand something called Kinetic Molecular Theory of Things. Kinetic Molecular Theory has many parts, but we'll introduce only a few here. One of the basic concepts of theory states that atoms and molecules have the energy of movement that we organize as temperatures. In other words, atoms and molecules are constantly moving, and we measure the energy of this movement as a material temperature. The more energy the material has, the more molecular movements will have, and the higher the temperature seen will be. One important thing that follows this is that the amount of energy that atoms and molecules have (and therein is the amount of movement) their interactions each other. Unlike simple billiard balls, many atoms and molecules are attracted to each other as a result of various intermolecular powers such as hydrogen bonds, van der Waals force, etc. Atoms and molecules that have a relatively small amount of energy (and movement) will interact strongly with each other, while those with relatively high energy will interact only a little, albeit at all, with others. Understanding Checkpoint Lowers material energy, __ How does this produce different conditions of things? Atoms with low energy interact strongly and tend to lock in places with respect to other atoms. Therefore, collectively, this atom forms a hard material, what we call solids. Atoms with high energy will move past each other freely, fly about the room, and form what we call gas. As it turns out, there are several known states on the matter; some of them detail below. Images © Corel Corporation Solids are formed when attractive force between individual molecules is greater than the energy that causes them to move separately. Individual molecules are locked in positions near each other, and cannot move past each other. Solid atoms or molecules remain in motion. However, the motion is limited to vibration energy; individual molecules remain in place and be on the other. As solid temperatures increase, vibration volume increases, but solids maintain its shape and volume as molecules are locked in relative places of each other. To see this example, click on the animation below that shows the molecular structure of the ice crystals. image © Corel Corporation Liquids form when energy (usually in the form of heat) system is increased and the rigid structure of the solid state is broken. In liquid, molecules can move past each other and meet other molecules; However, they remain close to each other like solids. Often in liquid, intermolecular power (such as hydrogen bonds indicated in the animation below) pull the molecules together and quickly break down. As liquid temperatures rise, the number of individual molecular movements increases. As a result, liquids can flow to take their container shape but they cannot be easily compressed because the molecules are already closed together. Therefore, the liquid has a shape that is not defined, but the amount is defined. In the animated example below, we see that liquid water consists of molecules that can move past each other, yet remain quite close to each other. images © Corel Corporation Gases form when the energy in the system outweighs all the interesting force between molecules. The gas molecules have little interaction with each other outside sometimes bruised bruises each other. In gas conditions, molecules move quickly and freely move in any direction, spreading long distances. As gas temperatures rise, the number of individual molecular movements increases. Gases grow to fill their containers and have low density. Because individual molecules are widely separated and can move easily in gas conditions, gases can be compressed easily and they have a non-subtle form. Solid, liquid, and gas are the most common states of things that exist on our planet. If you want to compare three states to each other, click on the comparison animation below. Note the difference in molecular movements in these three states. Interactive Animation:Image of Solid Liquid-Gas Comparison © NASA/JPL/Caltech Plasmas is a hot and ionted gas. Plasma is formed under very high energy conditions, so high, in fact, that molecules are riped apart and only free atoms exist. More surprisingly, plasma has so much energy that external electrons are actually ripped from individual atoms, thereof forming very energetic ion gas, charged. Because atoms in plasma exist as charged ions, plasma behaves differently from gases, thus representing the fourth state of things. Plasma can be seen only by looking upwards; High energy conditions that exist in stars such as individual atoms of our sun into plasma conditions. Comprehension Checkpoint __ has an unsigned and growing shape to fill their containers. As we have seen, increasing energy leads to a more molecular movement. On the other hand, reducing energy decisions in motion is less molecular. As a result, one prediction of the Kinetic Molecular Theory is that if we continue to reduce energy (measured as temperature) of the material, we will reach a point where all the molecular movements stop. The temperature in which the molecular movement stops is called absolute zero and has been calculated as -273.15 degrees Celsius. While scientists have cooled the ingredients for temperatures close to absolute zero, they never really reach absolute zero. The difficulty with observing the material at absolute zero is that to see the material, the light is needed, and the light itself transfers energy to the material, thus raising the temperature. Despite these challenges, scientists have recently observed the fifth state of things that only exist at temperatures very close to absolute zero. Bose-Einstein Condensates represents the fifth state of things only seen for the first time in 1995. The state is named after Satyendra Nath Bose and Albert Einstein who predicted its existence in the 1920s. B-E concentrates are gas superfluids that are cooled to very close zero zero temperatures. This strange situation, all condensate atoms reach the same quantum-mechanical state and can flow past each other without friction. More so B-E consensus can actually trap light, releasing it when the situation is damaged. Some other less common states of the matter have also been described or actually seen. Some of these states include liquid crystals, fermionic condensate, superfluids, supersolids, and things named aptly strange. To read more about this phase, see Phase in Our Resources for this module. The transformation of one state of matter to another state is called phase transition. A more common phase shift even has a name; For example, the term melting and freezing describes the transition of phase between solid and liquid conditions, and the terms of wipe and conspiration describes the transition between liquid and gas conditions. The transition of the phase occurs at a very accurate point, when energy (measured as temperature) of material in certain circumstances exceeds those allowed in the state. For example, liquid water can exist at various temperatures. Cold drinking water may be around 4ºC. Hot shower water has more energy and therefore probably around 40ºC. At this point, the energy introduced into the liquid will not go into increasing temperature; it will be used to send water molecules into gas conditions. Therefore, no matter how high the fire is above the kitchen, the boiling water pot will remain at 100ºC until all the water has suffered a transition to the gas phase. The excess energy introduced by high fire will accelerate the transition of liquid to gas; it will not change the temperature. The heat curve below reflects corresponding energy changes (indicated in calories) and water temperature as it experiences a phase shift between liquid and gas conditions. As can be seen in the graph above, as we move from left to right, the temperature of liquid water increases as energy (heat) is introduced. At 100ºC, the water begins to undergo a phase shift and the temperature remains continuous even though energy is added (the flat part of the graph). The energy introduced during this period goes towards breaking the intermolecular force so that individual water molecules can escape into gas conditions. Finally, when the transition is completed, if further energy is added to the system, the heat of gas water, or steam, will increase. This same process can be seen inversely if we only see the graph above starting on the right and moving left. Since the steam is cooled, the molecular movement of gas water and thus the temperature will decrease. When the gas reaches 100ºC, more energy will disappear from the system as an attractive force between the molecular redesign; However, the temperature remains continuous during the transition (the flat side of the graph). Finally, when the tapper is finished, the liquid temperature will begin fall as energy is withdrawn. Checkpoint Understanding As the heat is turned on under a boiling water pot, the temperature will __ until all the water has become gas. Phase transition is an important part of the whole world around us. For example, the energy is withdrawn when a sweat erodes from the surface of your skin allowing your body to control its temperature properly during hot days. Phase transition plays an important role in geology, affects the formation of minerals and possibly even earthquakes. And who can ignore the phase transition that occurs at about -3ºC, when the cream, perhaps with some strawberries or lumps of chocolate, begins to form solid ice cream? Now we understand what happens in a boiling water pot. The energy (heat) introduced at the bottom of the pot causes the transition of local phase liquid water to gas conditions. Because gas is less compact than liquid, this local phase shift forms a pocket (or bubble) of gases, which rise to the surface of the pot and burst. But nature is often more magical than our imagination. Although all we know about the transition state of things and phases, we still cannot predict where individual bubbles will form in boiling water pots. There are many states beyond solid, liquid, and gaseous, including plasma, condensate, superfluids, supersolids, and strange things. This module introduces the Kinetic Molecular Theory, which explains how atomic and molecular energy results in different conditions of things. The module also explains the process of transitioning the FASA in the matter. HS-C5.2, HS-PS1. A3, HS-PS1. A4, HS-PS2. B3 Anthony Carpi, Ph.D. Visionary Matter States Vol. CHE-3 (1), 2004. Top Page 12 Physical State and Property by Megan Cartwright, Ph.D., Anthony Carpi, Ph.D. Since 1924, The Macy's Thanksgiving Day Parade has wound up through 2.5 miles from New York City once a year. More than three million people gathered to enjoy the boisterous marching band, laughing at hundreds of clowns, and gawks in giant balloons floating over the march. Designed to look like cartoon characters like Snoopy (Fig. 1), each helium-filled balloon requires 90 operators to safely chew it through the march. Figure 1: Snoopy Balloons at the 2008 Macy's Thanksgiving Day Parade in New York City. Images © Ben W. (At first glance, the helium gas inside this balloon seems very different from the air outside of them. For one thing, balloons will be much less impressive if filled with air-instead of floating above the parade, they will be dragged along the ground. For others, while everyone enjoys the parade must breathe air to survive, they only have to breathe helium if they want a chequered voice. at different molecular, air and helium stages: Air is a mixture of nitrogen, nitrogen, and other gases, while helium is one gas. But helium and air have a lot of things in common with each other, and even with substances such as fatal carbon monoxide and flammable hydrogen. At standard temperatures and pressures, these materials are all gases, one of the common states of things (see our module of The Thing States for more information). All gases share the same physical properties. Like liquids, free gases flow to fill their containers. But even if the liquid has a defined amount, the gas does not have a defined volume or shape. And unlike liquids and solids, gases are very compressed. These general properties relate to the unique characteristics of gases: gas molecules are very far apart and rarely interact with each other. In solids, an interesting force and recover between molecules—the intermolecular team-so-strong they lock solid into a fixed form and size, as discussed in our Solid Properties module. In liquids, intermolecular power is weaker, and liquid molecules can move to each other. But the liquid molecules are still close enough that the intermolecular force affects nearby molecules (see the nature of our Liquid module). Gas molecules so far besides the intermolecular forces are ignored. Because gas molecules do not interact with each other, gases do not exist like different types such as liquids and solids do. Different types of liquids and solids (such as molecules and network solids) have properties that reflect the unique way their molecules interact. As a result, all gases share some common behavior. We can understand how any gas—whether helium or carbon monoxide—behaves by understanding laws governing gas behavior. Over the past four centuries, scientists have done many experiments to understand the usual behavior of gases. They have observed that the physical state of the gas—its condition—depends on the four variables: pressure (P), volume (V), temperature (T), and volume (n, in moles; see our module Mole: History and Use it for more information). The relationship between these variables is now known as gas law, which describes our current knowledge of how gases behave at macroscopic levels. But the relationship behind gas laws was unclear at first-they were unravelled by many scientists examining and testing their ideas about gas over the years. We now understand that air is a gas made of physical molecules (for more information, see our module on Atomic Theory). As these molecules move inside the container, they use force—known as pressure—on containers when they ricochet off their walls. Thanks to this behavior, we can expand cars, rubber shelves, and Macy's Day Parade balloons with gas. However, the idea that air is a material made of stressful molecules will be a strange idea scientists before the 17th century. Along with fire, water, and earth, the air is generally considered a fundamental ingredient, and not one that consists of other things. (For more information about this concept, see Our Initial Ideas on Things: From Democritus to Dalton module.) However, in 1644, Italian mathematician and physician Evangelista Torricelli proposed a strange idea. In a letter to fellow mathematicians, Torricelli described how he had filled a long glass tube full of mercry. When he seals one end and inverts the tube into the basin, only a few merquetries flow into the basin. The rest of the merzors live in a tube, filling a height of about 29 inches or 73.6 centimeters (Fig. 2). Torricelli suggested that it was the weight of pressing air on the mercry in the basin that forced the liquid into the tube (this is one of the first known devices that we now call barometers). Figure 2: Evangelista Torricelli experiments with mercury tubes and creates barometers. (Image from L'Atmosphere published in 1873.) Jesuit scientist Franciscus Linus has a different idea of what holds the merits in the tube. He suggested that mercury is being drawn by funiculus—an invisible substance that material materials come true to prevent vacuums forming between the top of the mercury tubes and sealed. British scientist Robert Boyle disagreed, and came up with an experiment to disable the idea of Linus' funiculus. Working with English physicist Robert Hooke, Boyle makes a curve-length glass tube like a cane, and he seals from a cane short leg. Resting the curves on the ground so that both ends point, Boyle pours in just enough mercury so that the silver liquid fills the curves and rises to the same height in each leg. The air was trapped this inside a sealed short leg. Boyle then poured in more mercury and observed with joy and satisfaction that the air trapped at the short end sealed supported a 29-inch high column (73.6 cm) of mercury in long legs—the same height that the mercury was achieved in the Torricelli toometer. However, since there is no cap on long legs, there may not be extra funiculus pulling extra merqueries. Boyle reasons that it must be a trapped air pressure (the so-called spring) that pushes the mercry up to 29 inches. To understand more about air pressure, Boyle poured more mercury into curve tubes. He recorded the height of the merquery column on the long leg, and the height of the air stuck on the short leg. After repeating these measures many times, Boyle was able to observe the relationship between the height of the air Growing height of merquery columns—indicators of pressure in the tube. Although scientists in Boyle's time generally didn't graph the data, we could see this relationship well Boyle data graph (Fig. 3). Figure 3: Robert Doyle's data plot that he recorded during his attempt at the mercry and stuck the air in a glass tube. Images © Krishnavedala Boyle's data show that when the air is squeezed into half its original volume, it doubles its pressure. In 1661, Boyle published his conclusion that the air volume was inversely related to its pressure. This observation of air behavior—and therefore, gas behavior—is a critical part of what we now call Boyle's law. Boyle's law states that as long as temperatures are kept continuous, total (V) fixed gas volume is propagation with pressure (P) (Fig. 4): Equation 1a Figure 4: Boyle Law states that as long as temperatures are kept constant, the amount of fixed gas is very propagational to the pressure Boyle Law can also be written as: 1b Equation For the amount of fixed gas at fixed temperatures, these constant , V_1) to (P_2, V_2), as volume decreases as pressure increases. Therefore, the P_1 x V_1 must be the same as the shy, and the P_2 x V_2 must also be the same as the slender. Because they're both similar to the same temper, gas and volume pressure under two different conditions is related like this: Equation 2 Back to helium balloons shaped like Snoopy, Boyle's law means that if you take balloons deep under the ocean, the poor Snoopy will mock because the pressure is very high and the helium will decrease significantly in volume. And if you take the balloon to the summit of Mount Everest, Snoopy will be bigger (and maybe even pop!) because atmospheric pressure is low and the helium will increase in volume. Which Entry Points Are two variables describing the gas situation are inversely related, according to Boyle's law? More than a century after Boyle's work, scientists have learned another important behavior of air: Air develops when heated, and hot air rises above cooler air. Taking advantage of this aerial behavior, the French brothers Joseph-Michel and Jacques-Étienne Montgolfier launched the first successful hot air balloon in Paris in 1783. The Montgolfiers balloon fascinated Jacques-Alexandre-César Charles, a self-taught French scientist interested in aeronautics. He had an idea of how to make better balloons. From its habit with contemporary chemical research, Charles knows that hydrogen is much lighter than air. In 1783, Charles built and launched the first hydrogen balloon (see Figure 4 for example balloon launches). Later that year, he became the first human to ride in a hydrogen balloon, which Almost 10,000 feet above Earth. Rajah 4: Jacques Charles and Nicolas Marie-Noel Robert stand in their hydrogen-filled waving flags, beginning their ascent Thousands of spectators were gathered in the foreground to witness the first gas balloon flight. Charles was very fortunate that he survived riding in a hydrogen balloon: On May 6, 1937, 36 people died when the Hindenburg plane, which smelled with flammable hydrogen, caught fire and crashed to the ground. The flammable hydrogen gas of the plane may have been erected by lightning bolts or sparks from static electricity, and the fire spread explosively across the ship within seconds. Although Charles never rode a balloon again, he remained blown away with gas inside the balloon. In 1787, Charles conducted an experiment comparing how balloons filled with different gases behaved when heated. Interestingly, she finds that balloons filled with different gases such as oxygen, hydrogen, and nitrogen extend the same amount when their temperature is heated from 0 to 80°C. However, Charles did not publish his findings. We just knew about his experiments as they were mentioned in the work of another French chemist and balloonist, Joseph-Louis Gay-Lussac. In 1802, Gay-Lussac published its results from the same experiment comparing nine different gases. Like Charles, Gay-Lussac concluded that it was common all-gas property to increase their totals the same amount when their temperatures were increased by the same degree. Gay-Lussac giving Charles credit for first observing this common gas behavior. The link between gas volume (V) and absolute temperatures (T, in Kelvin; to learn more about absolute temperatures, see our Temperature module) is now known as Charles's law. Charles's law states that when the pressure is kept continuous, a fixed amount of gas linearly increases its volume as temperatures rise (Fig. 5): Equation 3a Figure 5: Charles's law states that when the pressure is kept continuous, the amount of gas remains linearly increasing its volume as temperatures rise. Charles's law can also be understood as: 3b Equations For a fixed amount of gas on fixed pressure, these shyders will be the same, despite the amount of gas and temperature changes from (V_1, T_1) to (V_2, T_2). Therefore, the V_1/T_1 must be the same as the derived, and the V_2/T_2 must also be the same as the slender. As a result, the temperature and volume of gas under different conditions are related like this: Equation 4 This means that if we take Snoopy balloons to the North Pole, the balloons will shrink as the helium is cooled and decreased in volume. However, if we take balloons to a hot tropical island and helium temperatures rise, the helium will increase in volume, extending the balloons. Checkpoint Understanding When different gases are heated by the number of degrees equally, their amount will Be After his work on Charles's law, Gay-Lussac focused on figuring out how gas responded and combined. In 1808, he observed that many combine their volume in a simple and whole number ratio. While we understand now that the amount of gas is affiliated in the overall number ratio because that's how the gas molecules respond, Gay-Lussac does not suggest this explanation. This may be because of the idea of a combination of molecules the overall numbers have recently been proposed by John Dalton, who is a scientific rival of Gay-Lussac. (For further exploration of how gas molecules react, see our Chemical Equation module). It was Italian mathematician Amedeo Avogadro who realized that the idea of Dalton and Gay-Lussac complemented each other. Gay-Lussac's claim that the amount of gas combined in the full number ratio resembles Dalton's claim that atoms are combined in the overall numbers ratio to form a molecule. Avogadro reasons that the gas volume must be related to the number of its molecules. In 1811, Avogadro published his hypothesis that the same amount of gas had the same number of molecules. The Avogadro hypothesis has broken ground despite being largely ignored. Mathematicians rarely interact with other scientists, and he publishes his hypotheses with mathematical expressions that are not familiar to chemists. He also did not publish experimental data to support his hypotheses. It is 47 years before the Avogadro hypothesis will be widely recognized. In 1858, a former Avogadro student, Italian chemist Stanislao Cannizzaro, published an influential work on atomic theory. The work attracts Avogadro hypothesis and presents experimental data that supports hypotheses. Avogadro Law is based on the Avogadro hypothesis. Avogadro law states that at constant pressure and temperature, the volume of gas (V) is directly propagating with the number of molecules (n, in moles) (Fig. 6): Equation 5 Figure 6: Avogadro law states that on constant pressure and temperature, the number of gases is directly pronumlated with the number of molecules. We know that Snoopy balloons filled with helium will float above the parade, while similar balloons filled with air will be dragged along the ground. While helium and air differ in many ways, Avogadro's law means that if we compare the number of helium molecules and the number of air molecules needed to inflate the same Snoopy balloon, we will find that the numbers are the same. Checkpoint Understanding According to the laws of Avogadro, 1 liter of carbon gas toxic monoxide and 1 liter of flammable hydrogen gas both have the same: Because gas has common behavior described by gas law, we can understand and predict real gas behavior through the concept of ideal gas—theory of gas, the ideal ideal gas obtained from gas laws. This equation describes the relationship between all the modifiers examined in the gas law: pressure volume (V), volume (n, in moles), and absolute temperature (T, in Kelvins). Along with gas barking, R, these variables combine into ideal gas equations: Equation 6 Using the ideal gas equation, we can settle for any unknown variable, as long as we know the rest. The value for R depends on the units used for other variables (Table 2). Table 2: Constant gas values R for different units. Unit R's Value cal K-1 mol-1 1.9872 J K-1 mole-1 8.3145 L atm K-1 mole-1 0.0821 L Torr K-11 mol-1 62.364 Pa m3 K-1 mole-1 8.3145 The ideal gas law assumes that the ideal gas molecule has no volume, and does not suffer from intermolecular attraction or insulation. But real gas molecules have limited amounts and often have some (very small) interactions with each other. However, behavior and actual gas conditions are often predictable from the ideal gas equation, especially at standard temperatures and pressures. Under most circumstances the difference between real gas behavior and ideal gas behavior is very small so we can use the ideal gas equation for real gases. We will explore some of the circumstances in which real gas behaves differently from the ideal gas at the end of this module. Comprehension Checkpoint Ideal gas molecules have ideal gas laws also useful in situations where volume, n, gas are set, but pressure, volume, and temperature changes. Using ideal gas laws, we can associate the value of these three variables under different conditions. To do this, we need to first rearrange the ideal gas equation so that three variables that fluctuate the same nR: Equation 7a Equation 7b This relationship is referred to as affiliated gas laws. Because nR is a nrder, we can attrigue early (P_1, V_1, T_1) and the final state (P_2, V_2, T_2) gases like this: Equation 8 P 1 V 1 T 1 = P 2 V 2 T 2 From 1987 to 2012, the airbag (Fig. 6) saved nearly 37,000 lives in an airbag car saving lives because when the car stopped hard during the crash, the sensors triggered the chemical reaction Nitrogen gas infriends the airbags, which basically form a large pillow between the driver and the steering wheel. Pillows spread the force of the crash effect, helping to reduce the severity of the driver's injury. For airbags to work, it needs to expand its full nitrogen very fast—in 40 migrant collisions. For a 60-liter cylindric airbag to function properly, nitrogen gas needs to reach a pressure of 2.37 ATMs. At 25°C, how many nitrogen gas moles are needed to press airbags? We can think of this using the ideal gas equation. First, we list the values that and exchange them so that they have the same unit as the gas booster, R (0.0821 L-atm/mol-K). T = 25 ∘ C = (25 + 273) K = 298 K Next, we rearrange the ideal gas equation to solve mole, n: Finally, we finish the number of nitrogen gas moles to press airbags: n = (2.37 atms) × (60 L) (0.0821 L · ATM mol · K) × (298 K) Real gas often deviates from ideal gas when they get low especially when it is close to where gases are close to the gas. As the gas temperature decreases, its molecules move more slowly. This slower molecule is less able to overcome despite a weak intermolecular team in gas. This means that when the gas molecule will attack the walls of the container, the very small attraction it experiences for nearby gas molecules reduces its effects and the pressure given to the container. Therefore, actual gas at low temperatures puts lower pressure in the container (Fig. 7), compared to the ideal gas. Figure 7: Real gas at low temperatures puts lower pressure than predicted due to the attraction between gas molecules. Under high pressure, real gas often deviates from ideal gas because real gas molecules have volume, and attract each other. When the actual gas is under high pressure, its molecules are forced into smaller volumes. This smaller amount reduces the amount of free space the molecules need to move (Fig. 8). The amount of gas molecular space takes compared to the amount of space in the container—the number of molecular relatives—becomes larger. Figure 8: Under high pressure, actual gas has a larger amount than predicted due to the amount of molecules contained. This means that under real gas high pressure has a higher amount of ideal gas, since the ideal gas molecule has no volume. Furthermore, when real gas molecules are crowded close together, intermolecular forces can have more influence on molecular behavior. Interesting intermolecular forces draw molecules towards each other, which slow down the molecules and reduce their effects on container walls. Therefore, when it is under high pressure, real gases have slightly lower pressure than ideal gases. Checkpoint Understanding At low temperature, ideal gas in exerts __ containers of real gas. The characteristics of gas influence many important points, from the Earth's atmosphere to airbags to how we breathe. Among the breaths, the air pressure in our lungs is the same as the atmospheric pressure around us. When we in suck and use our rib cages and diaphrasm to develop our lung volume, air pressure decreases and external pressure forces the air in our lungs until stress is the same again—thus filling our lungs with the oxygen we need to survive. This module, we have focused on the usual properties of gases, and explored how these properties relate to a common set of behaviors called gas laws. We've also gotten ideal gas equations, and when these equations can—and cannot be used to predict real gas behavior. In other modules, we examine the properties of solid and liquid conditions of things, and explore molecular explanations for gas behavior with kinetic molecular theory. This module describes the nature of gas and explores how these properties relate to a common set of behaviors called gas laws. With a focus on Boyle Law, Charles Law, and Avogadro Law, an overview of 400 years of research shows the development of our understanding of gas behavior. This module presents an ideal gas equation and explains when these equations can—and cannot—be used to predict actual gas behavior. The Main Concept Unlike solid or liquid, the molecules in gases are very far apart and rarely interact with each other, which is why gases made of different molecules share the same behavior. Gas law describes the relationship between temperature, pressure, volume, and volume of gas. This law has been identified in experiments committed by various scientists for four centuries. Because gases share the same behavior, actual gas behavior at the given pressure (P), absolute temperature (T), volume (V), and volume (n, in moles) are often predictable by the ideal gas equation, PV = nRT, which perfectly describes the ideal gas behavior. Actual gas behavior deviates from ideal gases at very low temperatures and high pressure. Ashkenazi, G., James, S.G., & Jason, D.H. (2008). Similarities and differences in gas behavior: Interactive demonstrations. 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The original presentation of Boyle Law. Journal of Applied Physiology, 87(4): 1543-1545. -----. (2005). Robert Boyle's landmark book in 1660 with the first experiment on wired air. Journal of Applied Physiology, 98(1): 31-39. -----. (2014). Robert Hooke: early respiratory physiologist, polymath, and mechanical genius. Physiology (Bethesda), 29(4): 222-233. Whitaker, R.D. (1979). The initial development of kinetic theory. Journal of Chemical Education, 56(5): 315. Megan Cartwright, Ph.D., Anthony Carpi, Ph.D. Properties of Gases Visionlearning Vol. CHE-3 (9), 2016. Top Page 13 Physical and Real Estate country by Rachel Bernstein, Ph.D., Anthony Carpi, Ph.D. It's a classic prank: Fill saltshaker with sugar and wait for a meal to take an unexpected turn as your eating companions wonder why their cocks are oddly sweet. Or go in the other direction and you can really disturb someone's morning when they take the first salty sip of coffee. These pranks work very well because salt and sugar are almost indistinguishable by the naked eye: both are crystalline solids with the same structure. However, they have a very different flavor, and they behave differently as well. For example, you can pass electric current through salt water and light up light bulbs (you may have done this experiment yourself); but you can not do this with sugar water. Differences arise from different properties of both crystals, including the atoms that compose them and the true structure of the crystals themselves. In this module we will explore different types of solids and discuss how their structure relates to their behavior. Figure 1: Pile of salt (left) and sugar (right). From ancient Greece until the birth of modern chemistry in the 17th century, people might be confused about what made salt and sugar so different. Without today's tools to identify components in their crystals and structures, both of them will look similar to those as they do to our naked eyes today (see Figure 1 and 2). As scientists began to identify and characterize the elements in the 17th century 18th, they will be able to determine that salt is made of sodium Chlorine, while sugar consists of carbon, hydrogen, and oxygen, but they may still wonder how a combination of completely different elements leads to similar crystals. Figure 2: Close the view of salt (left) and sugar (right) crystals. Image © Salt: kevindooley; Sugar: Lauri Andler was not until the early 1900s that scientists could first see in crystals, when German scientist Max von Laue and father and English son of scientists William Bragg and Lawrence Bragg developed a method that used X-Rays to determine the microscopic structure of crystals. In fact, salt is the first solid investigated by this method, called X-ray crisstagraphy, which exposes the usual erosion of sodium and chlorine atoms. Applying X-ray crystagraphy to sugar reveals the same crystals but is not the same ordered well (Fig. 3). Similarities in their crystalline structural accounts for similarities in crystal appearance; However, the different types of atoms that make up each crystal and different arrangements of atomic accounts for behavioral differences between the two solids. X-ray cristography has also become a critical tool in modern biological research, helping to expose the structure of DNA double helix in the 1950s (see our DNA II: DNA module structure) and the structure of many simple and complex biological systems since that time. Now that researchers can see this level of detail through X-ray crisstagraphy and other methods, they can understand why some solids behave the way they do. And they can also use their understanding of the relationship between structure and behavior to design new and useful materials. Figure 3: Representative of salt atomic level (NaCl) and sugar (sukrosa, $C_{12}H_{22}O_{11}$). You may not think of salt and sugar as solid because when you see it in the kitchen they are small particles. But each of these particles is as solid as a wooden table, a glass window, or a piece of jewelry gold. Solids are a collection of atoms or molecules held together so that, under constant circumstances, they maintain defined shapes and sizes. Solids, of course, do not necessarily remain. Solid ice can melt to form liquid water at room temperature, and very high temperatures can be used to dilute solid iron so that it can be formed into skills, for example. Once the skill is formed and cools back to room temperature, though, its shape and size will not change on its own, as opposed to flat metals, which can be made to drip and alter shapes by gravity and mold. The same applies to ice cubes stored in the fridge: Once it is formed, their size and shape do not change. Solids have different shapes and sizes because it is formed when the force draws between individual atoms or molecules greater than the energy that causes them to move separately. Separate. Other words, atoms or molecules do not have enough energy to move and get stuck together in any form of them when they lose the energy to separate. (See our Article States module for more information on how solids differ from other states of things.) Salt and sugar are both crystalline solids. Other solid primary categories are called amorphous. Although crystal solids are ordered well at the atomic level, with each atom or molecule inhaccupting a certain point on erectile, amorphous solids triggered at atomic levels, with atoms or molecules held together in a fully random formation. Consider the inspector's game. Careful boards established with inspectors in each square are analogues to crystalline solids, while amorphous solids can be represented as pieces of inspectors randomly scattered across the board. Figure 4: Representation of glass atomic level (slika) and quarza. Quarza and glass are examples of the atomic levels of these two solid categories. Quartz is a crystalline solid that contains high slika (SiO_2) content. If we examine the structure of the quartz, we can see that the subunit sedlika is arranged accurately (see Figure 4). Glass, on the other hand, is amorphous solid. Although its smooth and transparent appearance might make it look like it must have a messy, organized microscopic structure, the opposite is true: Suction units are scattered across solids in a truly distractional fashion. Like quarza, glass has a very high content of slika (SiO_2). (See our Lion's Defining Minerals and Mineral Module for more on satay and quarza.) An important difference between crystals and amorphous solids is not what they are made of, but how they are made, and more precisely how their structures are sorted. The form of quartile on a very slow timescal scale, geology so atoms have time to achieve a highly ordered crystal structure, in which the atoms optimize the attractive force and minimize the repulsive force between them and therefore energize profitable. Glass, on the other hand, is made by melting sand (among other methods) and letting it cool quickly, freeze atoms in place, causing amorphous solid disorders. Amorphous solids often form when atoms and molecules are frozen before they have the opportunity to reach crystal order, which would otherwise be the structure of choice as it energized is favoured. One important consequence of an unregulated amorphous solid structure is that they do not always behave consistently or uniformly. For example, they can dilut various temperatures, in contrast to the very accurate solid crystal melting point. Return to the glass versus examples of quarza, the type of glass that frequently, called nuber lease soda, can dilute anywhere between and 1450°C, while cristobalite, polymorph quarza, melted right at 1713°C. In addition, amorphous solids break unpredictable and produce debris with irregular, often curved surfaces, while crystalline solids break along certain aircraft and at certain angles defined by crystal geometry. (See our Defining Minerals module for more information on how the external appearance of crystals reflects the fixed order of its atoms.) Checkpoint Understanding As an amorphous solid, glass has an accurate melting point. The crystalline structure determines more about solid than simply how it breaks. Structures are directly related to a number of important properties, including, for example, currisation and density, among others. To explain this relationship, we need to introduce four main types of crystalline solids - molecules, networks, ionics, and metals - which each described below. Individual molecules consist of atoms held together by strong covalent ties (see our Chemical Bonding module for more on covalent bonding). To form a solid molecule, these molecules are then arranged in certain patterns and held together by a relatively weak intermolecular team. Examples include ice (H_2O(s) – here means solid) and table sugar (sucrose, $C_{12}H_{22}O_{11}$). Water molecules and individual sugar each exist as their own independent entities that interact with their neighbors in a certain way to create a bessroots crystal solid. (See Figure 5). Figure 5: Two ice delegates: atomic-level organizations of molecules and ordinary ice cubes. Image © FDA (ice cubes) In the network solid, on the other hand, there are no individually defined molecules. The ongoing network of cohesive bonds holds together all atoms. For example, carbon can form two different network solids: diamonds and graphite. These ingredients consist of only carbon atoms arranged in two different ways. Diamond is a three-dimensional crystal that is the most difficult natural ingredient in the world. On the other hand, graphites are a solid two-dimensional network. Carbon atoms basically form a flat sheet, which is quite smooth and can slide past each other. Although these two substances are made of the same components very simple - only carbon atoms - their appearance and behavior are completely different because of the different types of bonding in solids. (See our Defining Minerals module for more information about diamonds and graphites.) The ability of one element to form multiple solids is called allergies. Figure 6: Diamond and graphite representations, including their atomic structure showing carbon atomic order. Images © Thatb Network can also incorporate a variety of elements. for example, consider quariza, the second most numerous ingredient in the Earth's crust. The chemical formulas for quariza are SiO_2, SiO_2, This formula shows a silicone-to-oxygen ratio and is not intended to imagine that there are different SiO_2 molecules now. Each silicon atom is tied to four different oxygen atoms and each oxygen atom is tied to two different silicon atoms, creating a large network of mutual bonds, as shown in Figure 6. (See Our Definition Minerals and Lion Mineral Module for more on quarza.) Ionic solids are similar to network solids in one way: there are no different molecules. But instead of atoms held together by leafy bonds, ionic solids consist of a positive and negative ion held together by ionic bonds. (See our Chemical Bonding module for more information on ionic bonding.) Table salt (sodium chloride, NaCl) is a typical ionic solid, as is anything called salt. Easy salt usually consists of one metal ion and one non-metal ion. In the case of sodium chloride, sodium is metal and chloride is not metallic. Salt can also consist of more complex ions, such as ammonium sulfate, whose components are ammonium (NH_4^+) and sulfates (SO_4^{2-}), are held individually together by convalued bonds and attracted to each other through ionic bonds. Finally, metal solids are kind of all their own. Although we discussed it lasting here, about three-quarters of the known elements are metal. You can read more about these metal elements in the Periodic Table of Elements modules. Here we will focus on how these elements behave as metal solids. Figure 7: This representation of the level of metal solid atoms indicates that electrons can easily move in solids. images © the Rafaelgarcia Metal atoms are held together by metal bonds, where atoms pack together and outdoor electrons can easily move in solids (Fig. 7). The metal bonds are meaningless, which means that metal atoms can remain tied while they roll into each other as long as some parts of their surfaces are in contact. The unique attributes of these metal bonds are largely responsible for some precious metal behaviors, including their circularity and malleability, which we discuss in the next section. The Center for Inspection of All Crystal Solids is held together by leafy bonds. As described in the previous section, crystal solids can vary in the composition of their atoms, bonds, and structures. Together, this attribute determines how different solids behave under different conditions. Solids have many different properties, including confusion, malleability, density, hardness, and optical delivery, to name a few. We will discuss some of these properties describes several ways that atoms and molecular structures drive function. As you read this lesson on your computer, you may not think about the way your computer uses to obtain the electrical power necessary to The wires are made of metal, perhaps copper, because metals generally have good electrical conductiveness. Electricity is basically an electron flow from one place to another, and in external electron metal bonds are quite free to move between adjacent atoms. This electron mobility means it is easy for the electric current to move from one end of the piece of metal to another. When electrons are introduced at one end of the wire by the electric current, this causes the electrons to move from one to another metal atom continuously down the wire, allowing the current to flow. In other solids, though, electrons are involved in covalent or ionic bonds and are therefore unable to run electricity, or do so just bad. Materials that do not run electricity are called electrical insulation. Heat, or heat, the viscincy is closely related to the softness of electricity. Just as metals are good electrical conductors, you probably know from the experience that they are also good at running heat as well. (That's why most kitchen pots, pans, and baking sheets are metal, so they can absorb heat from the kitchen or oven and hand them over to the food that is being cooked.) To understand how this works, consider that the temperature is a measure of how many moving molecules (see our conditions Matter and the Temperature module). For solids to run the heat, the movement of one molecule or atom should be easily transferred to its neighbor. The non-directional nature of the metal bonding makes this type of transfer quite simple, so the metal runs the heat well. In a solid network, on the other hand, where bonds are more rigid and the corners between atoms are strictly defined, such transfers are more difficult. Such solids are expected to have low thermal craftsweed and will be called thermal insulation. Figure 8: graphite sheets and carbon nanotube. Image © Graphite's NASA (nanotube) image is an interesting exception for this trend. Because of the certain energy and orientation of typical bonds in graphite sheets, they are quite good at running heat and electricity. You may have heard of carbon nanotubes, which are similar to graphite sheets but exist in the form of a tube (Fig. 8). This tube can run electricity and heat from one end to the other and is being tested for many possible applications, including in electrical circuits, solar cells, and textiles. Comprehension Checkpoint Metal runs heat and electricity well because bonds between atoms are Two additional properties, malleability and ductility, following trends similar to those for electrical and heat conductivity. Malleability describes the ability to hit solids into sheets without breaking them, and ductility refers to whether solids can be stretched to form wires. As you may have guessed, the metal tends to both malleable and ductile, ductile, due to the inconsistencies of metal bonds. On the other hand, covalent and ionic bonds, which are the direction and require certain geometry that results in a fixed three-dimensional lattice structure, make many other solid types fragile so that they break under force. Metal malleability and ductility are the important reason that metals are very useful. Their electrical confusion will be more useful if it is not possible to stretch them into wires that can then bend and shape at room temperature for an extraordinary array of applications. They also created some disadvantages though. Metal jewelry can be crushed and deformed at the bottom of the wallet, or a metal statue can be dented if it is dropped. Manufacturers must consider all the properties of the material they plan to work together to find the best option for each application. Another way to change solid is to melt it. Solid melting point depends on the strength of the interaction between its components: Stronger interaction means higher melting point. For molecular solids, melting means breaking down weak intermolecular forces (the force between different molecules), not the strong leafy bonds that hold individual molecules together, so compounds like sugar can easily melt on your kitchen. For network solids (held together by mutual bonds), ionic solids (held together by ionic bonds), and metal solids (held together by metal bonds), though, the melting temperature depends on the strength of certain bonds in each solid. Some metals have relatively low melting points, such as mercury, which in fact fluid at room temperature (its melting point is -38°C), while others, such as tungsten, melt only at very high temperatures (the melting point of tungsten is 3,422°C). Among the network solids, a quarza called tridymite melts at 1,670°C while the graphite melts at 4,489°C, and between ionic solids, sodium chloride melts at 801°C while lithium bromide melts at 552°C. Ionic bonds tend to be weaker than covalent and metal bonds, which is why the melting point for this salt is relatively low than most other examples of the melting point included here. Melting is one way to change solid shapes. Another approach is to dissolve solids into some type of liquid, in this case referred to as solvens. The extent of solid soluble in certain solvenuts is called its dissolution. Solids can be dissolved into different types of solvents, but for now we will focus on underwater disarray. Dissolving solids requires breaking different types of bonds for different types of solids. Dissolving metals requires breaking down metal bonds, and dissolving network solids requires breaking down leafy bonds. Both These bonds are very strong and difficult to break. Therefore, metals and solid networks generally do not dissolve in water. Water. the ring probably won't be so valuable if bands and stones are dissolved in the showers.) Instead, dissolving molecular solids requires breaking only a weak intermolecular team, not a bloody bond that actually holds individual molecules together. Therefore, molecular solids are quite soluble, since you may be able to guess given how we use sugar in so many drinks. Finally, to dissolve the ionic solids, the ionic bonds between atoms or molecules must be broken, which water is very good. Every atom or molecule in a solid ionic carries charges, and water molecules also carry charges due to polarity (see our Water: Properties and Behavior modules for more information). As a result, negative charges in water are attracted to positively charged ions, and positive charges in water are attracted to negative ions. This allows water molecules to dissolve ionic solids by separating parts, essentially trading lucrative ionic interactions in solid crystals with lucrative ionic interactions between individual ions and water molecules. Therefore, most salt is quite water-soluble. Both salt and sugar are quite dissolved in water, but because of the difference between ionic solids (salt) and molecular solids (sugar), salt water behaves differently from sugar water (remember the experimentation of light bulbs from the previous section). When salt dissolves in water, positive (Na^+) and negative (Cl^-) imposes a separate solid composing ion, creating a fermented particle fluid solution. These charged particles can take electrons and move them across solutions, effectively running electricity. When salt like ammonium sulfate is dissolved, the ionic tie-up between ions breaks, but the convalent bonds that hold individual complex ions together remain intact. By comparison, when sugar is decontaminated, each individual sucrose molecules separate from its neighbors but the sucrose molecules themselves remain intact and without charge, so they do not run electricity. The understanding of Checkpoint Dissolving the molecular solid requires fractional density, defined as the amount of mass that exists in a certain amount (see our Density module for more information), is another important property that depends on the structure and composition of the solid. It is important to note that although we attribute various types of crystalline solids as having certain structural characteristics, there are significant variations in each type as well. For example, metal solids don't all share the same atomic order. Atoms and molecules that make up crystals can wrap in many ways, affecting density (Fig. 9). Imagine a neatly ordered marble jar, with each dimple between marbles in a row filled with consecutive marble above. This leads to very high density. Gold takes approximately this type of packaging, causing a high density of 19.3 grams per cubic centimeter. Now imagine another jar where the marble is still neatly ordered, but each marble is sorted straight above the rest rather than pasted. This type of packaging leaves a much emptyer space in the jar because the dimples are not filled, so if the jar is the same size as the first jar, it cannot hold as much marble and less compact. Lithium, which is the least solid metal at 0.534 grams per cubic centimeter, is an example of this type of packaging. Figure 9: Two packaging geometry. The one on the left is a close arrangement, which produces high density; The one on the right is neaterly ordered, yet less wrapped and leaves more space, causing lower density. image © Vinícius Machado Vogt Another important variable is size. Larger marbles cannot be carefully packed like smaller marbles, although they are in the same order, so the contents of the jars will be less compact. However, if you are allowed to use marble of various sizes, you may be able to load small marble in the holes left between large marbles, which can lead to higher density than you would get from only small marbles. This principle is highly relevant for ionic solids, consisting of two different ions that usually vary in size. Bromide lithium, for example, is compact than pochium chloride. The difference in size between lithium and bromide is greater than the difference in size between phaloma and chloride, so lithium and bromide ions leave less free space when they pack together than phorum and chloride do, causing greater density. While solid properties may initially appear trivial, the unique characteristics of different solids affect almost every aspect of daily life in more ways than you think. Fine watches and, increasingly, other electronic devices use sapphire crystals instead of glass because strong network ties make sapphire very hard (in fact, it's the third hardest known material) and scratch-resistant. The common structure of ice molecules results in it being less compact than liquid water, and it can be said that without the life of this property on Earth will never exist. At less existent levels, this means that we can go ice skating in the frozen pool in winter even if it is not frozen all the way. Developing new solid materials with certain properties, such as electrical semiconductors and superconductors, is an active field of research with many potential applications. Solids are not the only substances with useful and entertaining properties, as we will see in the next module on liquids and gases. Solids are formed when teams hold atoms or molecules together stronger than the energy that moves them apart. This module demonstrates how multi-solid structures and compositions determine their properties, including conduction, humidity, density, and melting points. This module distinguishes two main categories of solids: crystals and amorphous. It then describes four types of crystalline solids: molecules, networks, ionics, and metals. See different solids explain how atomic and molecular structures drive functionality. The Main Solid Concept is a collection of atoms or molecules held together so that, under constant circumstances, they maintain defined shapes and sizes. There are two main categories of solids: crystals and amorphous. Crystal solids are ordered well at the atomic level, and amorphous solids erupt. There are four types of crystalline solids: molecular solids, network solids, ionic solids, and metal solids. The structure and composition of the solid atomic stage determines many of its macroscolic properties, including, for example, electrical and heat conductiveness, density, and dissolution. Rachel Bernstein, Ph.D., Anthony Carpi, Ph.D. Visionlearning Solid Properties Vol. CHE-3 (2), 2015. Top Page 14 Physical State and Property by Rachel Bernstein, Ph.D., Anthony Carpi, Ph.D. Water guesses out of the face. Honey oozes from squeeze bottles. Petrol flows out of the pump. These are only three examples of very diverse conditions: liquid. One of the main defining properties of the liquid is their ability to flow. Beyond this feature, though, the behavior of different liquids reaches variously. Some liquids flow relatively easily, such as water or oil, while others, such as honey or molasses, the flow is quite slow. Some are smooth, and some are sticky. Where does this different behavior originate? When it comes to interactions between different liquids, some mix well: Think Shirley Temple, made of ginger ale and grenades. Others, though, don't seem to be mixed at all. Consider the oil spill, where the oil floats in a sticky, iridescent layer on the water. You may also see a similar phenomenon in some separate salad dressings into layers of oil located above the layers of vinegar, which are mainly water. Why doesn't this liquid mix well? These varying behaviors arise mainly from the different types of intermolecular power found in liquids. In this module we will first discuss the liquid in the context of the other two main states of matter, solid and gas. Then we are going through a brief overview of the intermolecular team, and ultimately we will explore how the intermolecular power governs the way the liquid behaves. The liquid flows because the intermolecular forces between the molecules are weak enough to allow the molecules relative to each other. Intermolecular forces are troops between neighbours (This should not be confused with intramolecular forces, such as conciliation bonds and ionics, which are forces imposed in individual molecules to keep the atoms together.) The team drew when negative charges interacted with nearby positive charges and recovered when neighboring charges were the same, either positive or both negative. In liquids, intermolecular power can switch between molecules and allow them to move past each other and flow. (See Figure 1 for illustrations of various intermolecular forces and interactions.) Figure 1: Panel A shows a variety of interesting and recovered dipole-dipole interactions. Interesting interactions are shown in (a) and (b) with orientation where the positive ends are near the negative ends of other molecules. In (c) and (d), repulsive interactions are indicated by the orientation that juxtapose the positive or negative ends are polysized on the adjacent molecules. Panel B shows the liquid samples with some molecules both interesting and recovered with their policie-dipole interactions. Images © UC Davis ChemWiki Differ that with solid, where intermolecular forces are so powerful that they allow very little movement. Although molecules can squeeze in solids, they are basically locked into hardcore structures, as described in the properties of the Solid module. At the other end of the spectrum is gas, where molecules so far apart from that the intermolecular forces effectively do not exist and the molecules are completely free to move and flow freely. At the level of molecules, the liquid has some gas properties and some solids. First, the liquid shares the ability to flow with gas. Both the liquid and gas phases are liquid, which means that the intermolecular forces allow the molecules to move. In both phases, the ingredients have no fixed form and are otherwise formed by the container holding it. Solids are not liquid, but liquids share different important properties with them. The liquid and solid are both held together by strong intermolecular forces and far more compact than gases, leading to their description as the worst phase of things because both are somewhat incompressable. (Figure 2 shows the differences in gas, liquid, and solid at atomic levels.) Figure 2: Three states matter at atomic level: solid, liquid, and gas. images © Yelod Most materials can move between solid, liquid, and gas phases when the temperature is changed. Consider H20 molecules: It takes the form of ice, crystalline solids, below 0 ° C; water, liquid, between 0° and 100° C; and water vapor, or steam, gas, exceeds 100° C. This transition occurs because temperatures affect intermolecular attraction between molecules. When H20 is converted from liquid to gas, for instance, rising temperatures make molecules' increased energy until eventually outperforming intermolecular teams and molecules were able to move freely in the gas phase. However, the intramolecular forces holding H20 molecules together have not changed; H20 is still H20, regardless of its condition. You can read more about the phase shift in the State of Things module. Now that we have discussed how liquids are similar to and different from solids and gases, we can focus on the wide liquid world. First, though, we need to briefly introduce a wide variety of intermolecular powers that determine how fluids, and other states matter, behave. The Checkpoint Intermolecular Understanding team is As we earnact before, intermolecular forces are an interesting force or recover between molecules, differing from the intramolecular forces holding the molecules together. However, intramolecular forces play a role in determining the type of intermolecular power that can form. Intermolecular forces come in all sorts, but the whole idea is the same for all of them: Charges in one molecule interact with the charges in other molecules. Depending on the intramolecular forces that, such as pole confinement bonds or nonpolar convalued bonds, are present, charges can have multiple perseverances and strengths, allowing different types of intermolecular power. So, where do these charges come from? In some cases, molecules are held together by pole-confined bonds - meaning that electrons are not distributed equally between bonded atoms. (This type of bond is described in more detail in the Chemical Bonding module.) This uneverable distribution produces partial charges: Atoms with more electron affinity, that is, more electron atoms, have partial negative charges, and atoms with less electron affinity, less electron atoms, have partial positive charges. This unconscionable electron sharing is called a dipole. When two molecules with polyunsive covalent bonds are close to each other, they can form a profitable interaction if partial charges are partially aligned accordingly, as shown in Figure 3, forming a dipole-dipole interaction. Figure 3: In panel A, water molecules, H_2O, are indicated by unethical sharing of electrons causing partial negative charges around oxygen atoms and partial positive charges around hydrogen atoms. In panel B, three H_2O molecules interact well, forming a dipole-dipole interaction between partial charges. Hydrogen bonds are a very powerful type of dipole-dipole interaction. (Note that even if they are called bonds, they are not together or ionic bonds; they are powerful intermolecular power.) Hydrogen bonds occur when hydrogen atoms jointly to one of the few non-metals with high electronegativiti, including oxygen, nitrogen, and fluorine, embodies a strong dipole. Hydrogen hydrogen bonds hydrogen interactions of one of these molecules and atoms are more electroneregated in other molecules. Hydrogen bonds are present, and very important, in water, and are explained in more detail in our water: Properties and modules of Behavior. Hydrogen bonds and polished interactions require polar bonds, but another type of intermolecular power, called the London dispersing army, can form between any molecule, polar or not. The basic idea is that electrons in any molecule are always moving and sometimes, coincidentally, electrons can end up being circulated unevenly, manifesting a half-negative charge while on the molecular side with more electrons. These partial negative charges are offset by the same magnitud positive partial charge on the molecular side with less electrons, with positive charges coming from protons in the nucleus (Rajah 4). Temporary separa charges in these neighboring molecules can interact in the same way that polished remains interacting. The overall power of London's roasting powers depends on the size of the molecule: larger molecules can have larger temporary polishes, which leads to stronger London roasting powers. Rajah 4: Two nonpolar molecules with a sprinkling of sistetric molecules (panel A) can become poles (panel B) when the movement of electron rawak results in a temporary negative charge in one of the molecules, prompting an attractive (positive) charge elsewhere. Now, you might ask, if molecules can develop a temporary side-by-side charges that interact with each other, these temporary charges should also be able to interact with the eternally polished, right? And you'll be right. These interactions are called, very creative, polished-induced interactions. The accusation of some polar molecules interacts with electrons in nonpolar molecules and encourages them to move so that they are not circulated equally again, creating a dipole caused that can interact with the eternal forgetfulness of polar molecules (Rajah 5). Rajah 5: When polar molecules interact with electrons in nonpolar molecules (panel A), nonpolar molecules are encouraged to be polished and interact with polar molecules (panel B). As you may have pointed out, London's air force and ampole-induced interactions are weaker than polished interactions. These powers, as well as hydrogen bonds, are all van der Waals force, which is a general term for the attractive power between uncakaned molecules. There is more intermolecular power than what we have covered here, but with this brief introduction, we are prepared to return to the main event: the thaw, and how intermolecular power determines their nature and behavior. The Bush Point of Understanding Which interactions are stronger? If you have ever used oil to on the car, you know that it's nice and smooth. That's probably why you use it: it keeps fry pieces stirring from sticking to each other or pans, and it helps urinate the engine and other moving parts easily. One of the reasons good oils for this app is because they have low solidarity: liquid molecules do not interact especially strongly with each other because the intermolecular forces are weak. The main intermolecular forces present in most oils and many other organic liquids - liquids made mainly of carbon and hydrogen atoms, also referred to as non-polyunsitular fluids - are london dispersal teams, which for small molecules are the weakest types of intermolecular power. These weak forces lead to low solidarity. Molecules don't interact strongly with each other, so they can slide right past each other. At the other end of the unity spectrum, consider dewdrop on leaves in the early morning (Figure 6). How does such a thing exist if, as previously explained, the flow of fluids and taking the form of a container holding it? As described above and in the Water module, water molecules are held together by strong hydrogen bonds. These powerful forces lead to high solidarity: Water molecules interact with each other stronger than they interact with the air or leaves themselves. (Water interaction with leaves is an example of adhesion, or fluid interaction with something other than itself; we will discuss the adhesion in the next section.) Because of the high cohesion of water, molecules form a spherical shape to maximize their interactions with each other. Figure 6: Dew falls on leaves. the © of Cameron Whitman/iStockphoto This high Solidarity also creates surface tension. You may notice an insect walking on water in an outdoor pool (Figure 7), or seeing a small object like a paper located on the surface of the water instead of sinking; these are two examples of surface strains of water in action. Surface tension as a result of the powerful cubic forces of some liquids. These forces are strong enough to be maintained even if they suffer from outside powers such as insect gravity running across its surface. Figure 7: Strider water (Remigis Gerris), insects run normal water. Image © John Bush, MIT/NSF Adhesion is the tendency of compounds to interact with other compounds. (Remember that, on the other hand, unity is the tendency of compounds to interact with itself.) Adhesives help explain how fluids interact with their containers and with other liquids. An example of interaction with high adhesion is that between water and glass. Both water and glass are held together by bonds Therefore, both substances can also form a good pole interaction with each other, leading to high adhesions. You may have seen this interesting sticker force in action in the laboratory. Lab. water is in a glass graduating cylinder, for example, creepy water next to the glass, creating a concave curve at the top called meniscus, as indicated in the figure below. Water in a graduate cylinder made of several types of non-polar plastic, on the other hand, forms a flat meniscus because there is no attractive solid force or repellent between water and plastic. (See Figure 8 for comparison of passing poles instead of poles.) Figure 8: In the cylinder graduate A, made of glass, the meniscus is concave; in cylinder B, made of plastic, meniscus is flat. Image © Achim Prill/iStockphoto Comprehension Checkpoint When the intermolecular power is weak in the liquid, the liquid has low At the beginning of the module, we say that one of the defining characteristics of the liquid is their ability to flow. But among the liquid there is a huge variety in how easily this happens. Consider the ease with which you can pour yourself a glass of water, as opposed to the relative challenge of pouring thick motor oil and moving slowly into the engine. The difference is their viscracy, or resistance to flowing. Motor oil is quite viscous; water, not so much. But why? Before we delve into the difference between water and motor oil, let's compare water with other liquids: pentane (C_5H_{12}). Although we do not think of water as viscous, it is actually more viscous than pentane. Remember, water molecules form a strong bond of hydrogen to each other. Pentane, on the other hand, consists of only hydrogen and carbon atoms, notpolar, so the only type of intermolecular power it can form a relatively weak London disseminator team. Weaker intermolecular forces mean that molecules are easier to move past each other, or flow - therefore, lower viscracy. But both water and pentane are relatively small molecules. When we see a liquid made of larger molecules, its size is also played. For example, compare pentane to motor oil, which is a complex blend of large hydrokarbons far larger than a little pentane, and some with dozens or even hundreds of carbon in the chain. If you've ever poured motor oil into the engine, you know it's pretty clowed. Both liquids are nonpolar, and so have relatively weak intermolecular power; the difference is size. Large, bent on literal hydrokarbon motor oil can be trashed with their neighbors, which slows the flow. It's almost like a spaghetti pot: if you don't prepare it properly, you can end up with a custard noodle blob that's very hard to serve because they're all stuck together - in a sense, it's a blob of cigar pasta. More noodles - or smaller molecules – do not cile as much, so they tend to be less cific (Fig. 9). Figure 9: Group A consists of large molecules in cifectile blobs (bracity fluid) and Group B consists of smaller molecules with angles (less viscous liquid). Returning to the original comparison of motor oil compared to water, although water has strong intermolecular power, the larger molecular size in motor oil makes the oil more cloudy. There is another section for the story: temperature. Fluid heating makes it less grassy, since you may have observed if you have ever experienced how easy it is to pour maple syrup to your escort when the syrup has been heated than when it is cold. This is the case because the temperature affects both factors that determine the viscenity in the first place. First, increasing the temperature increases the kinetic energy of the molecules, which allows them to cope with the power of the intermolecules more easily. It also makes molecules move more, so large molecules that get turmoil when they are cold become more dynamic and more capable of skating past each other, allowing the liquid to flow more easily. Comprehension Checkpoint Motor oil pours slower than solver penetration because motor oil consists of When you think of water, you might think of its chemical formula, H_2O. This formula describes pure liquid consisting of only H_2O molecules, with absolutely no other components. The reality, though, is that the vast majority of the liquids we face are a complex mixture of many compounds. The solution is made of liquid solvenits where one or more dissolved are dissolved. Solutes can be solid, liquid, and gaseous. There are many, many common solutions that use water as solvenuts, including salt water and almost all kinds of flavored drinks. Carbon dioxide gas (CO_2) is a common gas milk in carbonated drinks, and tanol is a fluid that is celute in any alcoholic beverage. Although the solution is a mixture of multiple compounds, the properties discussed in the previous section still apply. Not all solutes dissolf in all solvens. You can dissolve a large number of a few dissolves in some liquid, and other dissolves are just a little dissolved in any solver. The basic explanation for reliability is that like a dissolution is like. Soluble non-soluble usually dissolves better in nonpolar liquids, and solutes dissolved better in polar fluids. For example, oil-based paint (and therefore nonpolar) requires non-pole solvents such as turpentine to clean; they will not dissolve in water, i.e. polarity. Table salt or sugar, on the other hand, both pole solids, are easily dissolved at high concentrations in water. More complex solutions include emulsions, colloids, and suspensions. In summary, emulsion is a well-spread mixture of two or more liquids that usually do not mix. Mayonnis, for example, is an oil emulsion, eggs, and vinegar or lemon juice, which is made by a very earnest mixing. Colloids and the hanging of both consist of unsolved zarahs in liquid. In colloids, colloids, particles that are not flammable are distributed in the liquid and will not separate. And suspension, on the other hand, is a liquid that contains larger particles that will eventually separate. Milk is a useful example of the difference between these two. Fresh milk is a suspension. It is a complex mixture of components that are usually not mixed - water, fats, proteins, carbohydrates, and more - and if left alone separate fat globules from the water-based part of the mixture. (Remember the separation of vinegar and oil in dressing salads? The process of milk separation is the same, with oily fat separating from the water.) Milk in most grocery stores, on the other hand, is colloid. The components do not separate thanks to a process called homogenization, which breaks the gllobules of fat into small enough particles that they can remain suspended in liquids. Checkpoint Understanding Which is a true statement about solutes? We have discussed many different liquids, with varying solidarity, adhesives, and viscosity, as well as other properties. But in addition to this variety, there are some substances that blur the difference between liquid and solid. For example, as a child you may have played with oobleck, a mixture of water and starch that got its name from Dr. Seuss's book. Oobleck is a slim material that can flow between your fingers if you hold it slowly in your hands but become hard and firm, almost solid, if you squeeze in. For more technical examples, consider the material used in LCD television displays and other electronic screens. LCD stands for Liquid Crystal Display. This does not mean that the display uses both liquids and crystals; this means that they use materials that are both liquid and crystal clear, at the same time. This may sound like a contradiction - crystals are solid, not liquid, you say - but those substances exist. The first liquid crystal discovered was a modified version of cholesterol, called cholesteryl benzoate. It is solid at room temperature and melts around 150°C, but then things get weird. At about 180°C, it changes phases again, but not from liquid to gas; it goes from clody liquid to cleaning the liquid. Austrian botanist and chemist Friedrich Reinitzer observed this extraordinary behavior in 1888 and discussed it with his colleague, German physicist Otto Lehmann. Lehmann then took over the investigation, studying cholesteryl benzoate and other compounds with similar two-melting behavior. When he saw a cloaky phase under his microscope, he discovered that the material appeared crystal, a solid defining feature. But that phase also flows, like liquid. In the year he coordinated the term liquid crystals to describe this phase, with properties between conventional liquid and crystalline solids. Liquid crystals play an important role in biology, especially in that should be liquid but must also maintain a normal structure. There are also some liquids that are so viscous you will not be blamed for thinking that they are solid, such as pitches, substances obtained from plants and petroleum. It seems almost solid, and shattered if hit by a hammer, but if left gravity it will flow very, very slowly. Several laboratories around the world conducted so-called pitch drop experiments, where they left several pitches in the funnel and waited to drip; about 10 years pass between each decrease (Fig. 10). Figure 10: Field Drop Experiment at the University of Queensland (battery shown for comparison of size). images of © John Mainstone & Amada44 Examples of materials behaving in a way that seems to oppose traditional definitions to phase out things that describe the complexity of science that exists and the natural world, although it comes to something that seems simple as determining whether material is fluid or solid. In this module, we have focused on determining and explaining the basic properties of the liquid, which provides the basis for you to think about the state of things in all their complexities. In other modules we discuss the solid and gas phase to help you distinguish the different physical properties in these states. When it comes to different liquids, some mix well while others do not; some pour quickly while others flow slowly. This module provides the basis for considering the situation of things in all their complexities. It describes the fundamental nature of the liquid, and explores how intermolecular power determines their behavior. The concept of unity, adhesives, and viscousness is defined. This module also examines how temperature and molecular size and type affect liquid properties. The Main Concept of Liquids shares some traits with solids - both considered the worst and somewhat unthinkable thing - and some with gas, such as their ability to flow and take the form of their containers. Some liquid properties, such as unity and adhesives, are influenced by the intermolecular power in the liquid itself. The viscity is influenced by both the intermolecular team and the size of the compound molecule. Most of the liquids we face in everyday life are actually solutions, solid mixtures, liquids or soluble gases in liquid solvenuts. HS-C6.2, HS-PS1. A3, HS-PS1. A4 Rachel Bernstein, Ph.D., Anthony Carpi, Ph.D. Visionlearning Liquid Property Vol. CHE-3 (5), 2015. Top