

Combined gas law worksheet

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The discovery of the electron, Nobel Prize-winning physicist Robert Millikan conducted an ingenious experiment that allowed for the specific value of the negative charge of the electron to be calculated. In his famous oil drop experiment, Millikan and co-workers sprayed tiny oil droplets from an atomizer into a sealed chamber (Millikan, 1911). The droplets were charged by a special corona discharge. The droplets were held in place by a voltage of 1000 V. By adjusting the voltage between the two electrical plates, Millikan applied an electrical force upward that exactly matched the gravitational force downward, thus suspending the drops motionless. When suspended, the electrical force and the force of gravity were working in opposite directions but were equal in magnitude. Hence, where q is the charge on the oil drop, E is the electric field, m is the mass of the oil drop, and g is the gravitational field. By measuring the mass of each oil drop and knowing both the gravitational and the electric field, the charge on each oil drop could be determined. Millikan found that there were differing charges on different oil drops. However, in each case the charges on the oil drops were found to be multiples of 1.60×10^{-19} coulombs. He concluded that the differing charges were due to different numbers of electrons, each having a negative charge of 1.60×10^{-19} coulombs, and hence the charge on the electron was found. Comprehension Checkpoint Millikan found that the different oil drops in his experiment Thomson's electron and Rutherford's nuclear model were tremendous advancements. The Japanese scientist Hantaro Nagaoka had previously rejected Thomson's Plum Pudding model on the grounds that opposing charges could not penetrate each other, and he counter-proposed a model of the atom that resembled the planet Saturn with rings of electrons revolving around a positive center. Upon hearing of Rutherford's work, he wrote to him in 1911 saying, "Congratulations on the simplicity of the apparatus you employ and the brilliant results you obtained." But, the planetary model was not perfect, and several inconsistent experimental observations meant much work was still to be done. At the time the electron was still thought of as a small particle, and it was thought to spin almost randomly around the nucleus of the atom. It would take the additional experiments and the genius of Neils Bohr, Max Planck, and others to make the paradigm shift from classical physics in which atoms consist of tiny particles and are governed by Newton's laws to quantum physics. The beginning of Quantum Theory, III: Wave-Particle Duality and the Electron, and IV: Quantum Numbers and Orbitals. Interactive Animation: Atomic and Ionic Structure of the first 12 elements The 19th and early 20th centuries saw great advances in our understanding of the atom. This module takes readers through experiments with cathode ray tubes that led to the discovery of the first subatomic particle: the electron. The module then describes Thomson's plum pudding model of the atom along with Rutherford's gold foil experiment that resulted in the nuclear model of the atom. Also explained is Millikan's oil drop experiment, which allowed him to determine an electron's charge. Readers will see how the work of many scientists was critical in this period of rapid development in atomic theory. Key Concepts Atoms are not dense spheres but consist of smaller particles including the negatively charged electron. The research on passing electrical currents through vacuum tubes by Faraday, Geissler, Crookes, and others laid the groundwork for discovery of the first subatomic particle. J.J. Thomson's observations of cathode rays provide the basis for the discovery of the electron. Rutherford, Geiger, and Marsden performed a series of gold foil experiments that indicated that atoms have small, dense, positively-charged centers - later named the nucleus. Millikan's oil drop experiment determines the fundamental charge on the electron as 1.60×10^{-19} coulombs. HS-C4.4, HS-C6.2, HS-PS1.A1, HS-PS1.A3 Faraday, M. (1838). VIII. Experimental researches in electricity. Thirteenth series. Philosophical Transactions of the Royal Society of London, 128: 125-168. Millikan, R.A. (1913). On the elementary electric charge and the Avogadro Constant. Physics Review, 2(2): 109-143. Rutherford, E. (1911). The scattering of α and β particles by matter and the structure of the atom. Philosophical Magazine, Series 6, 21(125): 669-688. Rutherford, E., & Natta, J.M. (1913). Scattering of α -particles by gases. Philosophical Magazine, Series 6, 26(154): 702-712. Rutherford, E. (1914). The scattering of α particles by matter and the structure of the atom. 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to its neighbor. The non-directional nature of the metallic bond makes this type of transfer relatively easy, so metals conduct heat well. In a network solid, on the other hand, the bonds are more rigid and the angles between the atoms are strictly defined, such transfer is more difficult. Such transfer should be expected to have a higher conductivity and would be called heat insulators. Figure 8: A graphite sheet and carbon nanotube. image © NASA (nanotube) Graphite is an interesting exception to this trend. Because of the specific energy and orientation of the typical bonds in graphite sheets, they are relatively good at conducting heat and electricity. You may have heard about SA carbon nanotubes, which are similar to graphite sheets but exist in the form of tubes (Figure 8). These tubes can conduct electricity and heat from one end to the other and are being tested for many possible applications, including in electrical circuits, solar cells, and textiles. Comprehension Checkpoint Metal conducts heat and electricity well because the bonds between atoms are Two additional properties, malleability and ductility, follow trends similar to those for electrical and thermal conductivity. Malleability describes the ability to hammer a solid into a sheet without breaking it, and ductility refers to whether a solid can be stretched to form a wire. As you may have guessed, metals tend to be both malleable and ductile, largely due to the non-directionality of metallic bonds. In contrast, covalent and ionic bonds, which are directional and require specific geometries resulting in fixed three-dimensional lattice structures, make many other types of solids brittle so they break under force. Metallic malleability and ductility are a crucial reason that metals are so useful. Their electrical conductivity would be much less useful if it weren't possible to stretch them into wires that could then be bent and shaped at room temperature for an incredible array of applications. They also create some drawbacks though. Metal jewelry can be crushed and deformed in the bottom of a purse, or a metal figurine can be dented if it's dropped. Manufacturers must consider all the properties of the materials they plan to work with to find the best option for each application. Another way to deform a solid is to melt it. A solid's melting point depends on the strength of the interactions between its components: Stronger interactions mean a higher melting point. For molecular solids, melting means breaking the weak intermolecular forces (the forces between different molecules), not the strong covalent bonds that hold the individual molecules together, so a compound like sugar can be easily melted on your stovetop. For network solids (held together by covalent bonds), ionic solids (held together by ionic bonds), and metallic solids (held together by metallic bonds) though, the melting temperature depends on the strength of the specific bonds in each solid. Some metals have relatively low melting points, like mercury, which is actually a liquid at room temperature (its melting point is -38°C), while others, such as tungsten, melt only at extremely high temperatures (tungsten's melting point is 3,422°C). Among network solids, a type of quartz called tridymite melts at 1,670°C while graphite melts at 4,489°C, and among ionic solids, sodium chloride melts at 801°C while lithium bromide melts at 552°C. Ionic bonds tend to be weaker than covalent and metallic bonds, which is why the melting points of these salts are somewhat lower than most of the other example melting points included here. Melting is one way of changing a solid's shape. Another approach is dissolving the solid into some type of liquid, in this case referred to as a solvent. The extent to which a solid dissolves in a particular solvent is called its solubility. Solids can be dissolved into a variety of types of solvents, but for now we will focus on solubility in water. Dissolving a solid requires breaking different types of bonds for different types of solids. Dissolving a metal requires breaking metallic bonds, and dissolving a network solid requires breaking covalent bonds. Both of these types of bonds are very strong and hard to break. Therefore, metals and network solids are generally not soluble in water. (Diamond rings probably wouldn't be as valuable if the band and the stone dissolved in the shower.) In contrast, dissolving a molecular solid requires breaking only weak intermolecular forces, not the covalent bonds that actually hold the individual molecules together. Therefore, molecular solids are relatively soluble, as you might have been able to guess given how we use sugar in so many drinks. Finally, to dissolve ionic solids, the ionic bonds between the atoms or molecules must be broken, which water does particularly well. Each atom or molecule within an ionic solid carries a charge, and water molecules also carry a charge due to polarity (see our Water: Properties and Behavior module for more information). As a result, the negative charges within water are attracted to the positively charged ions, and the positive charges within water are attracted to the negative ions. This allows the water molecules to dissolve ionic solids by separating the parts, essentially trading the favorable ionic interaction in the solid crystal with favorable ionic interactions between the individual ions and the water molecules. Therefore, most salts are relatively water-soluble. Both salt and sugar are quite soluble in water, but because of the differences between ionic solids (salt) and molecular solids (sugar), salt water behaves differently than sugar water (remember the light bulb experiment from the previous section). When salt dissolves in water, the positively (Na+) and negatively (Cl-) charged ions that compose the solid separate, creating a liquid solution of charged particles. These charged particles can pick up electrons and transfer them across the solution, effectively conducting electricity. When salts such as ammonium sulfate dissolve, the ionic bonds between the ions break, but the covalent bonds holding the individual complex ions together remain intact. By comparison, when sugar dissolves, each individual sucrose molecule separates from its neighbors but the sucrose molecules themselves remain intact and without charge, so they don't conduct electricity. Comprehension Checkpoint Dissolving a molecular solid requires breaking Density, defined as the amount of mass that exists in a certain volume (see our Density module for more information), is another important property that depends on the solid's structure and composition. It's important to note that although we described the different types of crystal solids as having certain structural characteristics, there is significant variation within each type as well. For example, metallic solids do not all share a similar arrangement of atoms. The atoms and molecules that make up crystals can pack in many different ways, which affects density (Figure 9). Imagine a jar of neatly ordered marbles, with each dimple between marbles in one row filled with a marble in the row above. This closely packed arrangement leads to a very high density. Gold takes on approximately this type of packing, resulting in its high density of 19.3 grams per cubic centimeter. Now imagine another jar where the marbles are still neatly ordered, but each marble is stacked directly on top of another instead of in the dimple. This type of packing leaves a lot more empty space in the jar because those dimples aren't filled, so if the jar is the same size as the first jar, it can't hold as many marbles and is less dense. Lithium, which is the least dense metal at 0.534 grams per cubic centimeter, is an example of this type of packing. Figure 9: Two packing geometries. The one on the left is a closely packed arrangement, which results in a high density; the one on the right is more neatly ordered, yet less packed and leaves more space, resulting in a lower density. image © Vincicus Machado Vogt Another important variable is size. Bigger marbles can't pack as closely as smaller marbles, even if they are in the same arrangement, so contents of the jar will be less dense. However, if you are allowed to use marbles of different sizes, you might be able to fit small marbles in the holes left between big marbles, which could lead to an even higher density than you would get from just the small marbles alone. This principle is particularly relevant for ionic solids, which are made up of two different ions that are usually different sizes. Lithium bromide, for example, is denser than potassium chloride. The size difference between lithium and bromide is greater than the size difference between potassium and chloride, so the lithium and bromide ions leave less empty space when they pack together than the potassium and chloride do, resulting in a higher density. While the properties of solids may at first appear trivial, the unique characteristics of different solids influence almost every aspect of daily life in more ways than you may think. Fine watches and, increasingly, other electronic devices use sapphire crystals instead of glass because the strong network bonding makes sapphire incredibly hard (in fact, it is the third hardest substance known) and scratch-resistant. The peculiar molecular structure of ice results in its being less dense than liquid water, and it can be argued that without this property life on Earth would have never have come into existence. On a less existential level, it means that we can go ice skating on frozen ponds in the winter even if it hasn't frozen all the way through. Developing new solid materials with specific properties, such as electrical semiconductors and superconductors, is an active area of research with many potential applications. But solids aren't the only substances with useful and entertaining properties, as we will see in the next modules on liquids and gases. Solids are formed when the forces holding atoms or molecules together are stronger than the energy moving them apart. This module shows how the structure and composition of various solids determine their properties, including conductivity, solubility, density, and melting point. The module distinguishes the two main categories of solids: crystalline and amorphous. It then describes the four types of crystalline solids: molecular, network, ionic, and metallic. A look at different solids makes clear how atomic and molecular structure drives function. Key Concepts A solid is a collection of atoms or molecules that are held together so that, under constant conditions, they maintain a defined shape and size. There are two main categories of solids: crystalline and amorphous. Crystalline solids are well ordered at the atomic level, and amorphous solids are disordered. There are four different types of crystalline solids: molecular solids, network solids, ionic solids, and metallic solids. A solid's atomic-level structure and composition determine many of its macroscopic properties, including, for example, electrical and heat conductivity, density, and solubility. Rachel Bernstein, Ph.D., Anthony Carpi, Ph.D. "Properties of Solids" Visionlearning Vol. CHE-3 (2), 2015. Top Page 19 Physical States and Properties by Rachel Bernstein, Ph.D., Anthony Carpi, Ph.D. Water gushes out of the faucet. Honey oozes out of a squeeze bottle. Gasoline flows out of the pump. These are just three examples of a highly diverse state of matter: liquids. One of the key defining properties of liquids is their ability to flow. Beyond this feature, though, the behaviors of different liquids span a broad range. Some liquids flow relatively easily, like water or oil, while others, like honey or molasses, flow quite slowly. Some are slippery, and some are sticky. Where do these different behaviors come from? When it comes to interactions between different liquids, some mix well: Think of a Shirley Temple, made of ginger ale and grenadine. Others, though, don't seem to mix at all. Consider oil spills, where the oil floats in a sticky, iridescent layer on top of the water. You may also notice a similar phenomenon in some salad dressings that separate into an oil layer that rests atop a layer of vinegar, which is primarily water. Why don't these liquids mix well? These varied behaviors arise primarily from the different types of intermolecular forces that are present in liquids. In this module we'll first discuss liquids in the context of the other two main states of matter, solids and gases. Then we will go through a brief overview of the intermolecular forces, and finally we'll explore how intermolecular forces govern the way that liquids behave. Liquids flow because the intermolecular forces between molecules are weak enough to allow the molecules to move around relative to one another. Intermolecular forces are the forces between neighboring molecules. (These are not to be confused with intramolecular forces, such as covalent and ionic bonds, which are the forces exerted within individual molecules to keep the atoms together.) The forces are attractive when a negative charge interacts with a nearby positive charge and repulsive when the neighboring charges are the same, either both positive or both negative. In liquids, the intermolecular forces can shift between molecules and allow them to move past one another and flow. (See Figure 1 for an illustration of the various intermolecular forces and interactions.) Figure 1: Panel A shows the variety of attractive and repulsive dipole-dipole interactions. Attractive interactions are shown in (a) and (b) with orientations where the positive end is near the negative end of another molecule. In (c) and (d), repulsive interactions are shown with orientations that juxtapose the positive or negative ends of the dipoles on adjacent molecules. Panel B shows a sample liquid with several molecules both attracting and repulsing with their dipole-dipole interactions. image © UC Davis ChemWiki Contrast that with a solid, in which the intermolecular forces are so strong that they allow very little movement. While molecules may vibrate in a solid, they are essentially locked into a rigid structure, as described in the Properties of Solids module. At the other end of the spectrum are gases, in which the molecules are so far apart that the intermolecular forces are effectively nonexistent and the molecules are completely free to move and flow independently. At a molecular level, liquids have some properties of gases and some of solids. First, liquids share the ability to flow with gases. Both liquid and gas phases are fluid, meaning that the intermolecular forces allow the molecules to move around. In both of these phases, the materials don't have fixed shapes and instead are shaped by the containers holding them. Solids are not fluid, but liquids share a different important property with them: liquids and solids are both held together by strong intermolecular forces and are much more dense than gases, leading to their description as "condensed matter" phases because they are both relatively incompressible. Figure 2 shows the differences of gases, liquids, and solids at the atomic level.) Figure 2: The three states of matter at the atomic level: solid, liquid, and gas. image © Yelod Most substances can move between the solid, liquid, and gas phases when the temperature is changed. Consider the molecule H2O: It takes the form of ice, a crystalline solid, below 0° C; water, a liquid, between 0° and 100° C; and water vapor, or steam, a gas, above 100° C. These transitions occur because temperature affects the intermolecular attraction between molecules. When H2O is converted from a liquid to gas, for instance, the rising temperature makes the molecules' kinetic energy increase such that it eventually overcomes the intermolecular forces and the molecules are able to move freely about in the gas phase. However, the intramolecular forces that hold the H2O molecule together are unchanged; H2O is still H2O, regardless of its state of matter. You can read more about phase transitions in the States of Matter module. Now that we've discussed how liquids are similar to and different from solids and gases, we can focus on the wide world of liquids. First, though, we need to briefly introduce the different types of intermolecular forces that dictate how liquids, and other states of matter, behave. Comprehension Checkpoint Intermolecular forces are As we described earlier, intermolecular forces are attractive or repulsive forces between molecules, distinct from the intramolecular forces that hold molecules together. Intramolecular forces do, however, play a role in determining the types of intermolecular forces that can form. Intermolecular forces come in a range of varieties, but the overall idea is the same for all of them: A charge within one molecule interacts with a charge in another molecule. Depending on which intramolecular forces, such as polar covalent bonds or nonpolar covalent bonds, are present, the charges can have varying permanence and strengths, allowing for different types of intermolecular forces. So, where do these charges come from? In some cases, molecules are held together by polar covalent bonds – which means that the electrons are not evenly distributed between the bonded atoms. (This type of bonding is described in more detail in the Chemical Bonding module.) This uneven distribution results in a partial charge: The atom with more electron affinity, that is, the more electronegative atom, has a partial negative charge, and the atom with less electron affinity, the less electronegative atom, has a partial positive charge. This uneven electron sharing is called a dipole. When two molecules with polar covalent bonds are near each other, they can form favorable interactions if the partial charges align appropriately, as shown in Figure 3, forming a dipole-dipole interaction. Figure 3: In panel A, a molecule of water, H2O, is shown with uneven electron sharing resulting in a partial negative charge around the oxygen atom and partial positive charges around the hydrogen atoms. In panel B, three H2O molecules interact favorably, forming a dipole-dipole interaction between the partial charges. Hydrogen bonds are a particularly strong type of dipole-dipole interaction. (Note that although they are called "bonds," they are not covalent or ionic bonds; they are a strong intermolecular force.) Hydrogen bonds occur when a hydrogen atom is covalently bonded to one of a few non-metals with high electronegativity, including oxygen, nitrogen, and fluorine, creating a strong dipole. The hydrogen bond is the interaction of the hydrogen from one of these molecules and the more electronegative atom in another molecule. Hydrogen bonds are present, and very important, in water, and are described in more detail in our Water: Properties and Behavior module. Hydrogen bonds and dipole-dipole interactions require polar bonds, but another type of intermolecular force, called London dispersion forces, can form between any molecules, polar or not. The basic idea is that the electrons in any molecule are constantly moving around and sometimes, just by chance, the electrons can end up distributed unequally, creating a temporary partial negative charge on the part of the molecule with more electrons. This partial negative charge is balanced by a partial positive charge of equal magnitude on the part of the molecule with fewer electrons, with the positive charge coming from the protons in the nucleus (Figure 4). These temporary partial charges in neighboring molecules can interact in much the same way that permanent dipoles interact. The overall strength of London dispersion forces depends on the size of the molon: larger molecules can have larger temporary dipoles, leading to stronger London dispersion forces. Figure 4: Two nonpolar molecules with symmetrical molecule distributions (panel A) can become polar (panel B) when the random movement of electrons results in temporary negative charge in one of the molecules, inducing an attractive (positive) charge in the other. Now, you might ask, if molecules can develop temporary partial charges that interact with each other, these temporary charges should also be able to interact with permanent dipoles, right? And you would be correct. These interactions are called, very creatively, dipole-induced dipole interactions. The partial charge of the polar molecule interacts with the electrons in the nonpolar molecule and "induces" them to move so they're not evenly distributed anymore, creating an induced dipole that can interact favorably with the polar molecule's permanent dipole (Figure 5). Figure 5: When a polar molecule interacts with the electrons in a nonpolar molecule (panel A), the nonpolar molecule is induced to become a dipole and interacts favorably with the polar molecule (panel B). As you might have guessed, London dispersion forces and dipole-induced dipole interactions are generally weaker than dipole-dipole interactions. These forces, as well as hydrogen bonds, are all van der Waals forces, which is a general term for attractive forces between uncharged molecules. There's a lot more to intermolecular forces than what we've covered here, but with this brief introduction, we're ready to get back to the main event: liquids, and how intermolecular forces determine their properties and behavior. Comprehension Checkpoint Which interactions are stronger? If you've ever used oil for cooking or working on a car, you know that it's nice and slippery. That's probably why you used it: It keeps stir-fry pieces from sticking to each other or the pan, and it helps engine pistons and other moving parts slide easily. One of the reasons oils are good for these applications is because they have low cohesion: the liquid molecules don't interact particularly strongly with each other because the intermolecular forces are weak. The molecules are held together by strong hydrogen bonds. These strong forces lead to high cohesion: the water molecules interact with each other more strongly than they interact with the air or the leaf itself. (The interaction of the water with the leaf is an example of adhesion, or the interaction of a liquid with something other than itself; we'll discuss adhesion in the next section.) Because of water's high cohesion, the molecules form a spherical shape to maximize their interactions with each other. Figure 6: Dew drops on a leaf. image © Cameron Whitman/Stockphoto This high cohesion also creates surface tension. You may have noticed insects walking on water on an outdoor pond (Figure 7), or seen a small object such as a paperclip resting on water's surface instead of sinking; these are two examples of water's surface tension in action. Surface tension results from the strong cohesive forces of some liquids. These forces are strong enough to be maintained even when they experience external forces like the weight of an insect walking across its surface. Figure 7: The water strider (Gerris remigis), a common water-walking insect. image © John Bush, MIT/NSF Adhesion is the tendency of a compound to interact with another compound. (Remember that, in contrast, cohesion is the tendency of a compound to interact with itself.) Adhesion helps explain how liquids interact with their containers and with other liquids. One example of an interaction with high adhesion is that between water and glass. Both water and glass are held together by polar bonds. Therefore, the two materials can also form favorable polar interactions with each other, leading to high adhesion. You may have even seen these attractive adhesive forces in action in lab. When water is in a glass graduated cylinder, for example, the water creeps up the sides of the glass, creating a concave curve at the top called a meniscus, as shown in the figure below. Water in graduated cylinders made out of some types of non-polar plastic, on the other hand, forms a flat meniscus because there are neither attractive nor repellant cohesive forces between the water and the plastic. (See Figure 8 for a comparison of polar and non-polar graduated cylinders.) Figure 8: In graduated cylinder A, made of glass, the meniscus is concave; in cylinder B, made of plastic, the meniscus is flat. image © Achim Prill/Stockphoto Comprehension Checkpoint When intermolecular forces are weak in a liquid, the liquid has low At the beginning of the module, we said that one of the defining features of liquids is their ability to flow. But among liquids there is a huge range in how easily this happens. Consider the ease with which you can pour yourself a glass of water, as compared to the relative challenge of pouring thick, slow-moving motor oil into an engine. The difference is their viscosity, or resistance to flow. Motor oil is quite viscous; water, not so much. But why? Before we dive into the differences between water and motor oil, let's compare water with another liquid: pentane (C5H12). While we don't think of water as viscous, it's actually more viscous than pentane. Remember, water molecules form strong hydrogen bonds with each other. Pentane, on the other hand, made up of just hydrogen and carbon atoms, is nonpolar, so the only types of intermolecular forces it can form are the relatively weak London dispersion forces. The weaker intermolecular forces mean that the molecules can more easily move past each other, or flow – hence, lower viscosity. But both water and pentane are relatively small molecules. When we're looking at liquids made up of bigger molecules, size comes into play as well. For example, compare pentane to motor oil, which is a complex mixture of large hydrocarbons much larger than little pentane, and some with dozens or even hundreds of carbons in a chain. If you've ever poured motor oil into an engine, you know it's pretty viscous. Both liquids are nonpolar, and so have relatively weak intermolecular forces; the difference is the size. The big, bendy motor oil hydrocarbons can literally get tangled with their neighbors, which slows the flow. It's almost like a pot of spaghetti: If you don't prepare it correctly, you can end up with a blob of tangled noodles that are very hard to serve because they're all stuck together – in a sense, it's a viscous pasta blob. Shorter noodles – or smaller molecules – don't tangle as much, so they tend to be less viscous (Figure 9). Figure 9: Group A consists of large molecules in a tangled blob (a viscous liquid) and Group B consists of smaller molecules with fewer entanglements (a less viscous liquid). Returning to our original comparison of motor oil versus water, even though water has such strong intermolecular forces, the much larger size of the molecules in the motor oil makes the oil more viscous. There's one more piece to the story: temperature. Warming a liquid makes it less viscous, as you may have observed if you've ever experienced how much easier it is to pour maple syrup onto your pancakes when the syrup has been warmed than when it is cold. This is the case because temperature affects both of the factors that determine viscosity in the first place. First, increasing the temperature increases the molecules' kinetic energy, which allows them to overcome the intermolecular forces more easily. It also makes the molecules move around more, so those big molecules that get tangled up when they were cold become more dynamic and are more able to slide past each other, allowing the liquid to flow more easily. Comprehension Checkpoint Motor oil pours more slowly than the solvent pentane because motor oil is made up of When you think of water, you might think of its chemical formula, H2O. This formula describes a pure liquid composed only of H2O molecules, with absolutely no other components. The reality, though, is that the vast majority of liquids we encounter are complex mixtures of many compounds. Solutions are made of a liquid solvent in which one or more solutes are dissolved. Solutes can be solids, liquids, and gases. There are many, many common solutions that use water as the solvent, including salt water and pretty much any type of flavored drink. Carbon dioxide (CO2) gas is a common gaseous solute in carbonated drinks, and ethanol is a liquid solute in any alcoholic drink. Although solutions are mixtures of multiple compounds, the properties discussed in the previous section still apply. Not all solutes dissolve in all solvents. You can dissolve some solutes in some liquids, and other solutes are only marginally soluble in any solvent. The underlying explanation for solubility is that "like dissolves like." Nonpolar solutes generally dissolve better in nonpolar liquids, and polar solutes dissolve better in polar liquids. For example, oil-based (and therefore nonpolar) paints require a non-polar solvent such as turpentine for clean up; they will not dissolve in water, which is polar. Table salt or sugar, on the other hand, both polar solids, easily dissolve at high concentrations in water. More complex solutions include emulsions, colloids, and suspensions. Briefly, an emulsion is a well-dispersed mixture of two or more liquids that don't normally mix. Mayonnaise, for example, is an emulsion of oil, egg yolk, and vinegar or lemon juice, which is made by very vigorous mixing. Colloids and suspensions both consist of insoluble particles in a liquid. In a colloid, the miniscule insoluble particles are distributed in a liquid and won't separate. And a suspension, on the other hand, is a liquid that contains larger insoluble particles that will eventually separate. Milk is a useful example of the difference between these two. Fresh milk is a suspension. It's a complex mixture of components that don't normally mix – water, fats, proteins, carbohydrates, and more – and if left alone the fat globules separate from the water-based portion of the mixture. (Remember the separation of vinegar and oil in salad dressing?) The milk separation process is similar, with the oily fat separating from the water.) The milk at most grocery stores, on the other hand, is a colloid. The components don't separate thanks to a process called homogenization, which breaks the fat globules into small enough particles that they can remain suspended in the liquid. Comprehension Checkpoint Which statement is true about solutes? We've discussed a lot of different liquids, with varying cohesion, adhesion, and viscosity, as well as other properties. But in addition to this already wide variety, there are some substances that blur the distinction between liquid and solid. For example, as a kid you may have played with oobleck, a mixture of water and starch that gets its name from a Dr. Seuss book. Oobleck is a slimy substance that can flow between your fingers if you hold it gently in your hands but becomes hard and firm, almost solid, if you squeeze it. For a more technical example, consider the material used in LCD television displays and other electronic screens. LCD stands for Liquid-Crystal Display. That doesn't mean that the displays use both liquids and crystals; it means that they use a material that is both liquid and crystal, at the same time. This might sound like a contradiction – crystals are solids, not liquids, you say – but such materials exist. The first liquid crystal discovered was a modified version of cholesterol, called cholesteryl benzoate. It's a solid at room temperature and melts at around 150°C, but then things get weird. At about 180°C, it changes phase again, but not from liquid to gas; it goes from cloudy liquid to clear liquid. Austrian botanist and chemist Friedrich Reintzer observed this unusual 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 double-melting behavior. When he looked at the cloudy phase under his microscope, he found that the material appeared crystalline, a defining feature of solids. But the phase also flowed, like a liquid. In 1904 he coined the term "liquid crystal" to describe this phase, with properties between those of a conventional liquid and crystalline solid. Liquid crystals play an important role in biology, particularly in membranes, which need to be fluid but also must retain a regular structure. There are also some liquids that are so viscous you wouldn't be blamed for thinking that they're solid, such as pitch, a substance derived from plants and petroleum. It appears almost solid, and shatters if hit with a hammer, but if left to gravity it will flow extremely, extremely slowly. A few labs around the world are running so-called pitch drop experiments, in which they leave some pitch in a funnel and wait for it to drip; about 10 years pass between each drop (Figure 10). Figure 10: The Pitch Drop Experiment at the University of Queensland (battery shown for size comparison). image © John Mainstone & Amada44 These examples of substances behaving in ways that seem to defy the traditional definitions for the phases of matter illustrate the inherent complexity of science and the natural world, even when it comes to something as seemingly simple as determining whether a substance is a liquid or a solid. In this module we have focused on defining and explaining the basic properties of liquids, which provides the foundation for you to think about states of matter in all their complexity. In other modules we discuss the solid and gas phases to help you contrast the different physical properties of these states. When it comes to different liquids, some mix well while others don't; some pour quickly while others flow slowly. This module provides a foundation for considering states of matter in all their complexity. It explains the basic properties of liquids, and explores how intermolecular forces determine their behavior. The concepts of cohesion, adhesion, and viscosity are defined. The module also examines how temperature and molecule size and type affect the properties of liquids. Key Concepts Liquids share some properties with solids – both are considered condensed matter and are relatively incompressible – and some with gases, such as their ability to flow and take the shape of their container. A number of properties of liquids, such as cohesion and adhesion, are influenced by the intermolecular forces within the liquid itself. Viscosity is influenced by both the intermolecular forces and molecular size of a compound. Most liquids we encounter in everyday life are in fact solutions, mixtures of a solid, liquid or gas solute within a liquid solvent. HS-C6.2, HS-PS1.A3, HS-PS1.A4 Rachel Bernstein, Ph.D., Anthony Carpi, Ph.D. "Properties of Liquids" Visionlearning Vol. CHE-3 (5), 2015. Top

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