

A glass of water looks simple enough to understand without much thought. It has no fixed shape, takes the form of its container, and can be poured from one place to another. Yet place that water in a freezer and it becomes a solid. Heat it sufficiently and it becomes a gas. Dissolve salt in it and the salt seems to disappear, although the water now tastes different and conducts electricity more readily. Let water remain in contact with iron and air, and rust can develop. The same familiar substance participates in several very different events. Chemistry begins when we stop treating those events as ordinary background activity and begin examining what the substances are made of, how their particles are arranged, and what changes when they interact. It is the study of matter—anything that has mass and occupies space—and the transformations matter can undergo. Its territory includes the food on a plate, the material in a phone battery, the dye in a shirt, the air entering a room, and the molecules that make living organisms possible.

The first useful distinction is between what we observe and what must be happening at a much smaller scale. We can watch sugar dissolve, but our eyes cannot follow individual sugar molecules separating and spreading through the water. We can see a copper surface change color, but the visible change alone does not explain the movement of electrons involved. Chemists connect these two scales: the large, observable world and the world of atoms, molecules, and ions. They also use a third language, symbols and equations, to represent what is happening. Much of the subject becomes confusing when someone presents the symbols before explaining the particles they represent. A formula such as H_2O is therefore best approached as a compact description of something real. It tells us that a water molecule contains two hydrogen atoms and one oxygen atom. The formula becomes useful once those words have meaning. Until then, it is merely a pattern to memorize.

An atom is the smallest unit of an element that retains that element's chemical identity. An *element* is a substance defined by the number of protons in its atoms. Hydrogen, oxygen, carbon, iron, gold, and helium are examples. At the center of an atom is its nucleus, a very small region containing positively charged protons and, in most atoms, electrically neutral neutrons. Negatively charged electrons occupy regions around the nucleus. A positive charge and a negative charge attract each other, while charges of the same sign repel. These electrical interactions are central to chemistry. Protons establish which element an atom belongs to, while electrons account for much of the way it interacts with other atoms. Hydrogen has one proton. Carbon has six. Oxygen has eight. If an atom has eight protons, it is oxygen, regardless of whether it is part of water, atmospheric oxygen, or a mineral. Changing the number of protons changes the element itself, which involves a nuclear transformation rather than an ordinary chemical reaction.

Pictures often show electrons circling a nucleus like planets traveling around a tiny sun. That image can introduce the idea of a central nucleus, but it becomes misleading if taken literally. Electrons do not follow the neat, visible tracks shown in those drawings. Chemistry describes them using *orbitals*, which are mathematical descriptions

associated with the regions where an electron is likely to be found. Their energies are also restricted to allowed states rather than any value whatsoever. This belongs to quantum mechanics, the branch of physics that describes matter and energy at very small scales. A beginner does not need to master its mathematics to learn chemistry, but one important consequence is worth understanding immediately: the arrangement of electrons is structured. Atoms do not carry electrons in an undifferentiated cloud where every electron has the same role. Some are held closer to the nucleus, while others occupy outer regions and are more directly involved in chemical bonding.

An electrically neutral atom has equal numbers of protons and electrons. The positive and negative charges balance. If an atom loses an electron, it has more positive charge than negative charge and becomes a positive ion, called a *cation*. If it gains an electron, it becomes a negative ion, called an *anion*. Losing an electron does not remove a proton, so the element's identity remains the same. A sodium atom and a sodium ion are both sodium because both have eleven protons, but they behave differently because their electron arrangements and charges differ. Neutrons introduce another distinction. Atoms of the same element can contain different numbers of neutrons; these versions are called *isotopes*. Carbon-12 and carbon-14 both have six protons, but they have different neutron counts. Some isotopes are stable, while others undergo radioactive decay. The existence of isotopes does not mean every atom is radioactive. It means that chemical identity and nuclear stability are related to different features of atomic structure.

The periodic table organizes the elements in increasing order of their atomic number, which is their number of protons. Its arrangement also reveals recurring patterns in electron structure and chemical behavior. The horizontal rows are called *periods*, and the vertical columns are called *groups*. Elements within a group often behave similarly because of similarities in their outer electrons. These outer electrons, called *valence electrons*, are especially important in bonding. In the main-group elements—the familiar groups on the left and right sides of the table—the group arrangement provides a useful guide to their number. Sodium and potassium, for example, belong to the same group and commonly form ions with a charge of positive one. Fluorine and chlorine belong to another group and commonly form ions with a charge of negative one. The table is therefore more than a catalog of names. It is a tool for anticipating behavior before conducting an experiment.

Those patterns should be treated as explanations to develop rather than rules to repeat without limits. Introductory chemistry often teaches that atoms “want eight electrons” in their outer shell. Atoms have no desires, and the so-called *octet rule* is a useful pattern for many common substances, not a universal command. Hydrogen is commonly stable with two electrons in its occupied shell, and numerous other substances do not fit the simple eight-electron picture. A more accurate starting point is that some electron arrangements are more energetically favorable than others. Atoms can become part of more stable arrangements by sharing electrons or by forming charged particles that attract one another. A *chemical bond* is the interaction that holds atoms or ions together in a substance. Understanding bonds helps explain why a substance has particular

properties and why replacing one substance with another can produce such a dramatic change.

In a covalent bond, atoms share electron density between them. A *molecule* is a discrete group of atoms held together by chemical bonds, most commonly covalent bonds. Oxygen gas consists mainly of O_2 molecules, each containing two oxygen atoms. Water consists of H_2O molecules, each containing two hydrogen atoms bonded to one oxygen atom. Oxygen gas is an element because only one type of atom is present. Water is a *compound* because it contains atoms of different elements chemically combined in a definite ratio. A molecule does not have to contain different elements, and a compound does not have to consist of separate molecules. These distinctions become easier when the terms are kept attached to their jobs: “element” identifies the kinds of atoms present, “compound” identifies a chemical combination of different elements, and “molecule” describes a particular kind of bonded particle.

Table salt illustrates why every compound should not be imagined as a collection of little molecules. Sodium chloride consists of positively charged sodium ions and negatively charged chloride ions arranged in an extended crystal lattice—a repeating, three-dimensional pattern. Their electrical attraction holds the solid together. Its formula, $NaCl$, indicates a one-to-one ratio of sodium ions to chloride ions. Chemists call this simplest ratio a *formula unit* rather than treating the crystal as separate $NaCl$ molecules. The properties of the compound differ greatly from those of the elements from which it can form. Sodium metal is highly reactive, and chlorine gas is hazardous, yet sodium chloride is the familiar substance used to season food. The explanation lies in chemical identity and structure: sodium ions in a salt crystal are not pieces of sodium metal, and chloride ions are not chlorine gas trapped inside a grain of salt. A compound must be understood as the substance it is, rather than as a collection of unchanged ingredients.

Metals have another useful bonding pattern. In a piece of copper or aluminum, some electrons are spread through the material rather than confined to a single bond between two particular atoms. This mobility helps explain why metals conduct electricity: electrical charge can move through them. Metallic bonding also helps explain why many metals can be bent, flattened into sheets, or drawn into wires without immediately crumbling. Their particles can shift relative to one another while the bonding remains broadly intact. Compare that with many ionic crystals, where shifting a layer can bring ions with the same charge close together. The resulting repulsion can cause the crystal to fracture. A copper wire and a salt crystal therefore behave differently under pressure for reasons connected to their internal structures. Chemistry gives the visible property a microscopic explanation.

Even knowing which atoms are present does not always tell us enough. Their arrangement can be decisive. Diamond and graphite are both forms of carbon, but their carbon atoms are connected differently. In diamond, each carbon atom is bonded into a strong three-dimensional network. In graphite, the atoms form sheets, with much weaker attractions between the sheets than the bonds within them. That structure allows the

layers to slide more readily, contributing to graphite's usefulness in pencil material and some lubricants. Different structural forms of the same element are called *allotropes*. Molecules can also contain the same numbers and kinds of atoms while arranging them differently; such substances are called *isomers*. The broader principle reaches throughout chemistry: composition tells us what is present, while structure tells us how it is organized. Both can determine the result. Having the same building materials does not guarantee having the same substance.

The states of matter provide another way to connect particle behavior with everyday observation. In a solid, particles maintain relatively fixed positions, although they still vibrate. In a liquid, particles remain close together but can move past one another. In a gas, particles are much farther apart on average and move throughout the available space. Ice, liquid water, and water vapor contain the same kind of molecule, H_2O . Their differences arise from movement and arrangement, including the attractions between neighboring molecules. Melting ice does not ordinarily break water molecules into hydrogen and oxygen. It changes how the molecules are organized and how freely they move. This is a *physical change*: the substance's chemical identity remains the same. By contrast, a *chemical change* produces different substances through changes in bonding or electron arrangements.

The attractions between water molecules are especially important. Within each molecule, oxygen pulls the shared electrons toward itself more strongly than hydrogen does. This ability to attract shared electrons is called *electronegativity*. Water also has a bent molecular shape, so its uneven distribution of charge does not cancel out. The oxygen side is partially negative, and the hydrogen sides are partially positive. Such a molecule is described as *polar*. Nearby water molecules attract one another in a particularly significant interaction called hydrogen bonding. Despite the name, a hydrogen bond between water molecules is different from the covalent oxygen–hydrogen bonds inside each molecule. The distinction explains why water can evaporate while remaining water. Enough energy allows some molecules to overcome the attractions holding them near their neighbors, without tearing apart the molecules themselves.

Water's structure also helps explain an unusual and important property: ordinary ice is less dense than liquid water. *Density* is mass divided by volume, a measure of how much matter is packed into a given amount of space. As water freezes under ordinary conditions, hydrogen bonding organizes the molecules into a more open structure than they have in the liquid. The same mass therefore occupies more volume, lowering its density and allowing ice to float. Mass and weight should be distinguished here as well. Mass describes a property of matter measured in units such as grams and kilograms; weight is the gravitational force acting on that mass. A sample taken to the Moon would have less weight because lunar gravity is weaker, but its mass would remain essentially the same. Chemistry usually measures quantities of material through mass because it provides a consistent basis for comparing samples.

Most materials encountered outside a laboratory are mixtures. A *mixture* contains two or more substances together without requiring them to form a new compound in a fixed ratio. Air is a mixture of gases. Saltwater is a mixture of water and dissolved salt. Soil contains a complicated assortment of minerals, organic material, water, and gases. In a *homogeneous mixture*, the composition is uniform at the scale being considered; a well-mixed salt solution is an example. A *heterogeneous mixture* has distinguishable regions or components, as in oil and water after they separate. Mixtures can often be separated using physical properties. Filtration separates certain suspended solids from a fluid. Distillation uses differences in volatility—the tendency to enter the gas phase—to separate substances. A filter that catches sand, however, will not ordinarily remove dissolved salt. The particles involved and their relationship with the water are different.

A solution is a homogeneous mixture in which one or more substances are dispersed in another. The dissolving medium is called the *solvent*, and a dissolved substance is called a *solute*. When sugar dissolves in water, water molecules interact with sugar molecules and help separate them from the crystal. The sugar molecules remain sugar molecules as they spread through the solution. Sodium chloride behaves differently: the crystal separates into sodium and chloride ions, which become surrounded by water molecules. This is why saltwater can conduct electricity through the movement of dissolved ions, while a solution of pure sugar does not provide the same supply of charge-carrying particles. Dissolving therefore does not mean vanishing, and it does not always involve exactly the same microscopic process. The substance remains present even when the eye can no longer distinguish it from the surrounding liquid.

The phrase “like dissolves like” is often used to introduce solubility, the extent to which a substance can dissolve under particular conditions. Polar solvents tend to dissolve many polar substances and ionic compounds, while nonpolar solvents often dissolve nonpolar substances more readily. This is a useful starting point rather than a guarantee. Water does not dissolve every salt well, and many details of molecular structure matter. Oil and water separate because the interactions available when they mix do not favor a uniform mixture under ordinary conditions. Soap helps because its molecules have regions that interact differently with water and oily material. In water, suitable soap molecules can gather into structures called *micelles*, with their water-avoiding portions directed inward and their water-interacting portions facing outward. Oily material can be incorporated into these structures and carried away during rinsing. Cleaning becomes easier to understand when it is viewed as a problem of molecular interactions.

How much solute is present also matters. *Concentration* expresses the amount of a substance relative to an amount of solution or solvent, depending on the measurement being used. Two cups can both contain saltwater while having very different concentrations. If one contains twice as much salt in the same final volume of solution, it is more concentrated. Adding water lowers the concentration through *dilution*, although it does not remove the salt already present. This distinction is useful far beyond a classroom: “less concentrated” does not mean “less total substance” unless the quantities being compared support that conclusion. A small container of concentrated

material and a large container of dilute material could contain the same total amount of solute. Whenever a label gives a percentage or a concentration, the accompanying definition matters. A percentage by mass and a percentage by volume describe different comparisons.

Chemical reactions occur when substances are transformed into other substances. The starting substances are called *reactants*, and the substances formed are called *products*. During ordinary chemical reactions, atoms are rearranged rather than created or destroyed. This leads to conservation of mass: in a closed system, where matter cannot enter or leave, the total mass remains the same through a chemical reaction. A burning piece of wood appears to lose matter if only the remaining ash is considered, but much of its material has become gases and airborne particles. Oxygen from the surrounding air has also participated. Accounting for only the visible solid produces an incomplete comparison. Chemistry trains us to define the *system*, meaning the portion of the world being studied, and to track what crosses its boundary. A disappearing solid or an expanding cloud does not suspend conservation; it makes careful accounting necessary.

Chemical equations record that accounting. Consider the complete combustion of methane, a major component of natural gas: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$. In words, one methane molecule reacts with two oxygen molecules to produce one carbon dioxide molecule and two water molecules. The small numbers within the formulas, called *subscripts*, belong to the identities of the substances. CH_4 contains four hydrogen atoms per molecule. The larger numbers placed before formulas, called *coefficients*, specify relative quantities. The equation is balanced because both sides contain one carbon atom, four hydrogen atoms, and four oxygen atoms. Balancing an equation means adjusting coefficients to preserve the number of each kind of atom. Changing a subscript instead would change the substance itself. H_2O and H_2O_2 are water and hydrogen peroxide, respectively; the extra oxygen atom does not merely mean “more water.”

The ratios in an equation become practical through a counting unit called the *mole*. Individual atoms and molecules are too small to count one at a time in ordinary laboratory work. A mole is a specified number of entities, just as a dozen means twelve, except that the mole contains approximately 6.022×10^{23} entities. The notation means 6.022 multiplied by ten raised to the twenty-third power, an enormous quantity suited to extremely small particles. One mole of water contains that many water molecules. One mole of oxygen gas contains that many O_2 molecules and therefore twice as many oxygen atoms. A mole is a count, not a particular mass: a mole of heavy particles weighs more than a mole of lighter particles, just as a dozen bowling balls weighs more than a dozen tennis balls. The size of the count stays the same while the mass depends on what is being counted.

Molar mass connects that particle count to a measurable mass. It is the mass of one mole of a substance, commonly expressed in grams per mole. Water’s molar mass is approximately 18 grams per mole because its two hydrogen atoms and one oxygen atom contribute their respective atomic masses. About 18 grams of water therefore contains

one mole of water molecules, while about 36 grams contains two moles. This allows a chemist to weigh material and calculate how many reacting particles it represents. *Stoichiometry* is the calculation of quantities using the ratios in balanced chemical equations. The word looks intimidating, but the central idea resembles preparing meals from a recipe. If every sandwich requires two slices of bread and one slice of cheese, the available quantities determine the maximum number of sandwiches. Chemical reactions follow particle ratios with far greater precision, but the counting logic is approachable.

The ingredient that runs out first is called the *limiting reactant*. Suppose the balanced equation requires two moles of oxygen for each mole of methane. Having one mole of methane and only one mole of oxygen does not provide enough oxygen for the complete reaction as written. Under that simplified calculation, oxygen limits how much methane can react by that pathway. The other reactant may remain in excess. The greatest product quantity predicted from the available reactants is the *theoretical yield*. The amount actually obtained is the *actual yield*. Their comparison gives *percent yield*, calculated by dividing actual yield by theoretical yield and multiplying by 100. Actual yield may be lower because a reaction does not go to completion, competing reactions occur, or material is lost during separation and handling. The equation provides a quantitative expectation; the experiment reveals how closely real conditions approach it.

Energy requires equally careful accounting. An especially common misunderstanding is that breaking a chemical bond releases energy. Breaking a bond requires energy; forming a bond releases energy. Whether an overall reaction releases or absorbs energy depends on the complete balance of the bonds and interactions being broken and formed. In an *exothermic reaction*, energy is transferred from the reacting system to its surroundings, often as heat. In an *endothermic reaction*, the system absorbs energy from its surroundings. A fuel can release energy overall because the formation of its reaction products releases more energy than is required to disrupt the starting arrangements. The fuel's usefulness lies in the difference between the starting and ending states. Describing energy as "stored in bonds" can be convenient, but it becomes misleading if it suggests that simply snapping bonds is what supplies the energy.

Temperature and heat are related but different concepts. Temperature reflects the thermal state of a material and is associated with the motion and energy distribution of its particles. Heat is energy transferred because of a temperature difference. A large bathtub of warm water can transfer more energy as it cools than a tiny cup of hotter water, because the amount of material matters alongside its temperature. Substances also differ in *specific heat capacity*, the energy needed to raise the temperature of a given mass by a specified amount. Water has a relatively high specific heat capacity, which helps explain why large bodies of water warm and cool gradually. Energy added during a phase change may alter particle arrangement without raising temperature. At a fixed pressure, pure water undergoing steady boiling can absorb energy while its temperature remains approximately constant; the energy is being used to convert liquid into vapor.

A reaction that can release energy does not necessarily begin rapidly on its own. It may need to overcome an *activation energy*, an energy barrier between reactants and products. A useful comparison is a ball resting behind a low ridge on a hillside. The ball could end up lower down, but it first needs enough energy to get over the ridge. The comparison is imperfect, yet it captures why a favorable overall change can require an initial push. Reaction speed belongs to *chemical kinetics*, the study of how fast reactions occur and the steps through which they proceed. Temperature, concentration, available surface area, and the reaction mechanism can all affect that speed. Increasing temperature often speeds reactions because more particles can overcome the relevant energy barrier. Increasing concentration can increase the frequency of useful encounters between reacting particles. Neither effect should be treated as an unlimited rule independent of the substances involved.

A *catalyst* speeds a reaction by providing a pathway with a lower activation barrier. It participates in the mechanism but is regenerated rather than consumed overall. It does not create energy, and it does not change the overall energy difference between reactants and products. A catalyst can therefore make an available reaction proceed more quickly without turning an energetically unfavorable process into a favorable one by itself. *Enzymes* are biological catalysts, usually proteins, that make many of the reactions required for life proceed at useful rates under suitable conditions. Their shapes and chemical features help them interact with particular reacting substances. This explains why changes that disrupt an enzyme's structure can interfere with its function. In living systems, chemistry depends on organization as well as ingredients: having the appropriate molecules present is not sufficient if the structures that coordinate their reactions are damaged or absent.

Some reactions are reversible, meaning products can react to regenerate the original reactants. In a closed system under suitable conditions, such a reaction may reach *dynamic equilibrium*. At equilibrium, the forward and reverse reactions continue, but they occur at equal rates, so the overall composition remains constant. Equal rates do not mean equal quantities. There could be much more product than reactant while the two opposing processes proceed at the same speed. Imagine people entering and leaving a room at equal rates: the number inside stays steady even though movement continues. Changing concentration, pressure, or temperature can alter the equilibrium composition, depending on the reaction. A catalyst helps the system approach equilibrium faster by accelerating both directions, but it does not change the equilibrium composition at a given temperature. This distinction separates the question "How quickly does change happen?" from "What composition does the system approach?"

Acids and bases introduce another important family of reactions. In one widely used definition, an *acid* donates a proton, while a *base* accepts one. In this setting, "proton" commonly refers to H^+ , a hydrogen ion. In water, that proton is associated with surrounding water molecules rather than existing as a completely isolated particle. It is often represented as hydronium, H_3O^+ . Acidity is commonly described using *pH*, a logarithmic measure related to hydrogen-ion activity. For a beginner working with dilute

solutions, it is useful to understand it as a scale connected to the effective availability of hydrogen ions. At about 25°C, neutral water has a pH near 7; lower values are acidic and higher values are basic. Because the scale is logarithmic, a one-unit decrease corresponds to a tenfold increase in hydrogen-ion activity. A solution at pH 3 therefore has about one hundred times the hydrogen-ion activity of one at pH 5, rather than being merely “two units more acidic” in an ordinary linear sense.

Acid strength and acid concentration must be kept separate. A *strong acid* transfers its proton to water essentially completely under the usual introductory conditions, while a *weak acid* does so only partly, establishing an equilibrium. Concentration describes how much acid is present in a given amount of solution. A weak acid can be concentrated, and a strong acid can be dilute. The everyday meaning of “weak” can therefore mislead someone into assuming a substance is automatically harmless. Safety depends on more than that classification. Acids and bases can react through *neutralization*, often producing water and an ionic compound called a salt, although the resulting solution is not necessarily exactly pH 7. A *buffer* is a system that resists large pH changes when limited amounts of acid or base are added. It works because its components can react with added hydrogen ions or with substances that would otherwise remove them. Buffers have finite capacity; they cannot absorb unlimited chemical change.

Electron transfer produces another broad category of reactions. *Oxidation* is the loss of electrons, and *reduction* is the gain of electrons. The two occur together because electrons lost by one species must be accepted elsewhere. Their combined name is *redox*. Although the word oxidation developed through reactions involving oxygen, oxygen does not have to participate in every redox reaction. Rusting involves oxidation of iron as part of a larger process involving its surroundings. Batteries use redox chemistry in a deliberately organized way. By separating the places where electrons are released and accepted and providing an external conducting path, a battery can drive electrons through a circuit. Ions move within the battery to maintain charge balance. The device therefore turns a chemical energy difference into electrical work. Recharging a rechargeable battery uses supplied electrical energy to drive chemical changes back toward a state that can discharge again.

Carbon opens the door to an enormous portion of chemistry because it can form strong bonds with itself and with many other elements. It commonly forms four covalent bonds, allowing chains, branches, rings, and more complex structures. *Organic chemistry* studies carbon-based compounds, with certain carbon-containing substances, such as carbon dioxide and many carbonates, conventionally treated as inorganic. In this scientific usage, “organic” does not mean pesticide-free, healthy, natural, or produced on a particular kind of farm. A synthetic medicine and a molecule in a leaf can both belong to organic chemistry. Much of a molecule’s behavior depends on its *functional groups*, recognizable arrangements of atoms that contribute characteristic chemical properties. An oxygen–hydrogen group attached in one structural setting behaves differently from that same general group in another. Learning these patterns gradually allows a chemist to look at a molecular structure and anticipate how parts of it may interact.

Large molecules extend those patterns further. A *polymer* consists of many smaller units connected into long chains or networks. The smaller building units are called *monomers*. Plastics include synthetic polymers, while cellulose in plants and proteins in living organisms demonstrate that large chain-based molecules are also fundamental to nature. A protein is built from amino-acid units, and its sequence influences how it folds and what it can do. Two materials described simply as “polymers” may therefore have very different properties. Chain flexibility, the strength of attractions between chains, the presence of cross-links, and the identities of the repeating units all matter. Stretchy rubber, a rigid plastic container, and plant fiber cannot be understood through size alone. Their structures determine how they respond to heat, force, solvents, and other conditions.

Chemical learning also requires distinguishing an observation from an explanation. Bubbles show that gas is present or being released, but they do not by themselves prove that a chemical reaction is producing a new gas. Boiling water bubbles because liquid water is changing into vapor. A color change may indicate a reaction, but it can also result from mixing colored materials or changing particle size. A solid appearing in a solution may be a *precipitate*, an insoluble substance formed through a reaction, but crystallization can also occur when a dissolved substance comes out of solution as conditions change. Chemists investigate by controlling variables, taking measurements, comparing alternatives, and testing whether an explanation accounts for all the evidence. An instrument reading also has uncertainty, meaning a range within which the measured value can reasonably lie. Careful science includes that limitation instead of presenting every digit as exact.

A strong beginning in chemistry comes from repeatedly connecting the same few ideas in new settings. A grain of salt becomes an ionic lattice. Water becomes a polar molecular substance whose attractions influence dissolving, freezing, and heat storage. A reaction equation becomes a count of atoms and a guide to measurable quantities. A battery becomes organized electron transfer. A cooking process becomes a combination of heat movement, physical changes, and chemical reactions whose rates depend on conditions. None of those explanations eliminates the familiar experience; each makes it more understandable. The next time sugar dissolves in a glass, its disappearance need not be the end of the observation. There are still molecules present, still measurable mass, still interactions determining where those molecules can go. The visible glass has not changed very much. What has changed is how much of its activity the observer can recognize.