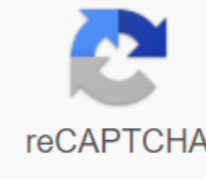




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## Atomic theory test pdf

Atoms are made of protons, neutrons and electrons. The periodic table organizes the elements depending on their atomic size and other properties. If you see this message, it means that we are having trouble downloading external resources on our site. If you're behind a web filter, please make sure the domains no.kastatic.org and no.kasandbox.org unlocked. 4th, 5th, 6th, 7th, 8th, 9th, 10th, 11th, 12th Year ZO Not We have not found any results on the atomic%20theory%20test. Please check your spelling and try again. Atomic Theory and Structure by Adrian Dingle, B.Sc., Anthony Carpi, Ph.D. Top Page 2 Atomic Theory and Structure Adrian Dingle, B.Sc., Anthony Carpi, Ph.D., This is an updated version of our module atomic theory I. For the previous version, please go here. By the late 1800s, John Dalton's view of atoms as the smallest particles that make up all matter was influenced for about 100 years, but this idea is about to be challenged. Several scientists working on atomic models have found that atoms are not the smallest particles that make up matter, and that different parts of the atom have very clear characteristics. The English scientist Michael Faraday can reasonably be considered one of the greatest minds ever in the field of electrochemi and electromagnetism. Somewhat paradoxically, all of Faraday's pioneering work was carried out before the discovery of the fundamental particle on which these electrical phenomena depend. However, one of Faraday's first experimental observations was a decisive precursor to the discovery of the first subatomic particle, the electron. As early as the mid-17th century, scientists experimented with glass tubes filled with the then-so-famous rarefied air. Rare air refers to a system in which most gaseous atoms were removed, but where the vacuum was not complete. In 1638, Faraday noted that when the current passed through such a pipe, there was an arc of electricity. The arc began on a negative plate (known as a cathode) and passed through a tube to the opposite-charged anode (Faraday, 1838). In his experiments, Faraday observed luminescence, which began part of the way down the tube, and traveled to the anode. This left the area between the cathode and the luminescence at the beginning of the luminescence, which was not illuminated, and subsequently became known as the dark space of Faraday (Figure 1). Faraday was unable to fully explain his observations, and it took several further developments in terms of pipe technology before a deeper understanding emerged. Figure 1: Discharge glow in a low pressure tube caused by an electric shock. Like what Faraday saw, the tube shows the dark space between the glowing around the cathode (left, negatively charged) and the anode (right, positively charged). image of © Comprehension Checkpoint Faraday иэвєstєn иэвєstєn detection and the name of the electron. In 1857, German glassblower Heinrich Geissler, working for fellow countryman and physicist Julius Plucker at the University of Bonn, improved the quality of the vacuum that could be achieved in such tubes. However, Geissler's pipes still contained enough gas atoms that when the electric current traveled in the tube, there was an interaction between them, causing the pipes to glow. In the mid-19th century, these Geissler pipes were largely nothing more than curiosity, but interesting, curiosity, which turned out to be a precursor to neon lights. Englishman William Crookes repeated experiments similar to the experiments of Faraday and Geisler, but this time with new and improved vacuums. The number of gas atoms (and therefore pressure) has been drastically reduced in Crooks pipes. This had an interesting effect: Faraday's dark space was observed even further down the pipe, again extending from the cathode to the anode. In addition to expanding the dark space, fluorescence was observed on the glass behind the anode at the positive end of the tube. When further experiments showed that the shadows of objects placed in a tube were thrown at the glass behind the anode, the German physicist Johann Hittorf suggested that the shadows should have been created by something traveling in a straight line from the cathode to the anode. Another German physicist Eugene Goldstein christened these invisible rays cathode rays. As discussed in our module Early Ideas about Matter, Dalton's atomic theory suggested that the atom was indivisible, i.e. that it was the smallest particle that made up matter, and that all matter was based on this single unit. Experiments with cathode beam tubes dramatically changed this view when they led to the discovery of the first subatomic particle, J.J. Thomson is an English physicist who worked with cathode beam tubes similar to those used by Crooks and others in the mid-19th century. Thomson's experiments (Thomson, 1897) went further than before him, and provided evidence of the properties of something that Hittorf hinted at. Thomson noted that the cathode rays were deflected by magnetic fields and that the deviation was the same regardless of the source of the rays. This suggests that the rays are versatile in their properties and that they have some magnetic charge. Thomson also demonstrated that the cathode rays are charged because they can deflect by an electric field. He found that the rays veered toward the positive plate and away from the negative plate, thus determining that they were made of negatively charged particles of some kind. Finally, Thomson simultaneously applied both electric and magnetic fields to the cathode beam. Knowing the strength of the fields used and measuring deflection in the tube, it was for the first time able to measure the speed of particles in the stream. Then, by measuring the deviation of the beam, changing the two fields, Thomson was able to measure the mass-to-charge ratio of the particles in the stream, and he found something amazing. Negative particles had a mass-to-charge ratio that was more than 1,000 times lower than that of a hydrogen atom, suggesting that the particles were incredibly tiny - much smaller than the smallest atom known. This fact allowed Thomson to definitively say that the atom was not a fundamental building block of matter and that there were smaller (subatomic) particles. Thomson initially called these particles calves, but later they became known as electrons. The JJ Thomson checkpoint found that the cathode rays were made from Thomson's discovery made sense of all previous observations made by Faraday, Geisler, and Crooks. Stretching a tube filled with gas, but partially in a vacuum, the electrons eventually burst into these gas atoms, knocking down some of their electrons and making them fluorescent. The dark space that Faraday first noted was due to the distance required for electrons to accelerate to the speed required to ionize tube gas atoms. In the best vacuum achieved in Crooks tubes, electrons could travel further distances without interacting with gas molecules due to the lower density of molecules in the tube, thereby expanding the dark space. Understanding the checkpoint The less gas was in the cathode beam tube, the dark space could be observed. When the electron was discovered, Thomson proposed a brand new model of the atom, which was known as the Plum Pudding Model. The model was so-called, as it mimics the British desert with the same name as dried fruit (primarily raisins not plums), scattered in the body of the suta and the eggs that made the dough. In his model, Thomson suggested that negatively charged electrons (similar to raisins) were randomly distributed between what he called a sphere of uniform positive electrification (similar to the dough or the pudding body) (see Figure 2). Figure 2: Thomson's plum pudding is a model of an atom, showing a positively charged sphere containing many negatively charged electrons in a random arrangement. Thomson's model of the atom as a loose cluster of positive and negative particles persisted until 1911, when Ernest Rutherford, a former Thomson student, advanced atomic theory for another cutout. In 1908-1911 Ernest Marsden and Hans Geiger conducted a series of experiments led by Ernest Rutherford at the University of Manchester in England. In these experiments, alpha particles (tiny, positively charged particles) were released into a thin piece of gold foil (Figure 3). Under Thomson's Plum The atom model, the sphere of uniform positive electrification, was considered so scattered that tiny, fast-moving alpha particles would pass straight through it. Similarly, the electrons in the model were thought to be so tiny that any electrostatic interactions between them and positive alpha particles would be minimal, so the path of alpha particles would hardly be affected. Figure 3: Gold foil experiment developed by Rutherford, Marsden and Geiger. A beam of positively charged alpha particles was shot in a piece of gold foil. The screen around the foil recorded the effects of alpha particles. As predicted, Rutherford and his colleagues noticed that most alpha particles pass directly through the golden foil, and some particles deviate at small angles. However, contradicting what the plum pudding model predicted, a few bounced at very sharp angles, some even fly straight back to the source! These particles acted as if they were colliding with a solid object, like a tennis ball bouncing off a brick wall (Figure 4). Figure 4: In an experiment with gold foil, Rutherford and his colleagues expected to see alpha particles passing through mostly empty plum pudding-style atoms. However, they noticed that alpha particles sometimes ricocheted at an acute angle, indicating that there was something more solid in the atom than previously thought. The fact that most alpha particles passed directly through the gold foil showed Rutherford that atoms are mostly made up of empty space. However, unlike the Thomson Plum Pudding model, Rutherford's work showed that the atom had a dense, positively charged area, which caused the observed repulsion and rollback of alpha particles. Rutherford was impressed by these observations and famously said: It was absolutely the most incredible event ever happened to me in my life. It was almost as incredible as if you shot a 15-inch shell on a piece of paper and it came back and hit you. On reviewing I realized that this scattering backwards must be the result of a single collision, and when I made the calculations, I saw that it was impossible to get anything of that magnitude unless you took a system in which most of the mass of the atom was concentrated per minute of the nucleus. That's when I came up with the idea of an atom with a minute-long massive center carrying a charge. During a series of experiments and works (Rutherford, 1911, 1913, 1914) Rutherford developed a model of an atom with a dense, positively charged area of the atom in the center, now known as the nucleus - and a nuclear model of the atom was born. Understanding Checkpoint Rutherford and his colleagues were surprised that after the discovery of the electron, Nobel Prize-winning physicist Robert Millikan held a genius that allowed us to calculate the specific value of the negative electron charge. In their famous experiment with drops of oil, Millikan and his colleagues sprayed tiny drops of oil from the sprayer into a sealed chamber (Millikan, 1913). Oil droplets fell down, under the influence of gravity, into the space between two electric plates. There they are charged, interacting with the air that was ionized by X-rays. Figure 5: Millikan's oil drop experiment in which he observed drops of oil fall between two electrical plates where droplets became ionized by X-rays. By regulating the voltage between the two electric plates, Millikan applied an electrical force upwards that matched the gravitational force downwards, thus suspending the droplets motionless. When suspending, the electric force and gravity worked in opposite directions, but were equal in size. Hence: where q is a charge on the fall of oil, E is an electric field, m - the mass of oil droplet and g - gravitational field. By measuring the mass of each oil drop and knowing both the gravitational and electric field, it is possible to determine the charge on each drop. Millikan found that there were different charges in different drops of oil. However, in each case, the charges of oil drops were found to be multiples of 1.60 x 10-19 pendants. He concluded that different charges were caused by a different number of electrons, each of which has a negative charge of 1.60 x 10-19 pendants, and therefore a charge on the electron was found. At the Millikan checkpoint, he found that the various droplets of oil in Thomson's electron and Rutherford's nuclear model were huge advances. Japanese scientist Hantaro Nagaoaka had previously rejected Thomson's plum model on the grounds that opposing charges could not penetrate each other, and he proposed a model of an atom resembling the planet Saturn with rings of electrons orbiting a positive center. Upon hearing of Rutherford's work, he wrote to him in 1911. Congratulations on the simplicity of the apparatus you use and the brilliant results you have received. However, the planetary model was not ideal, and several contradictory experimental observations meant that much more needed to be done. At the time, the electron was still considered a small particle, and it was thought to rotate almost randomly around the nucleus of the atom. It will take additional experiments and genius of Neil Bohr, Max Planck and others to make a paradigm transition from classical physics, in which atoms are composed of tiny particles and governed by the laws of motion, to quantum mechanics, in which electrons behave like waves and exhibit strange and exotic behavior. (See our interactive animation by comparing orbital and quantum models of the first 12 elements.) To find out strange behavior of quantum physics, read other records in our series of atomic theory. II: Bor and the beginning of quantum theory. III: Wave-particle duality and electron, and IV: quantum numbers and orbits. Interactive animation: The atomic and ionic structure of the first 12 elements of the 19th and early 20th centuries saw great progress in our understanding of the atom. This module takes readers through experiments with cathode beam tubes, which led to the discovery of the first subatomic particle: the electron. The module then describes the Thomson Atom plum pudding model along with Rutherford's experiment with gold foil, which led to a nuclear atom model. Also explained is Millikan's experiment on oil drop, which allowed him to determine the charge of the electron. Readers will see how the work of many scientists was critical during this period of rapid development in atomic theory. The key concepts of atoms are not dense spheres, but are composed of small particles, including a negatively charged electron. Studies on the passage of electric current through vacuum tubes Faraday, Geisler, Crooks and others laid the groundwork for the discovery of the first subatomic particle, J.J. Thomson's cathode observations are the basis for the electron discovery. Rutherford, Geiger and Marsden conducted a series of gold foil experiments that showed that atoms have small, dense, positively charged centers, later named the nucleus. Millikan's oil drop experiment (Einstein, 1905). Sometimes described as Einstein's Annus Mirabilis year), the documents combined, and combined with Planck's work, allowed Bohr to marry the nature of atoms with physics to usher in a new dawn of understanding in atomic theory. In 1913, relying on Planck and Einstein's quantitative assessment theories, Bohr suggested that the electron itself orbits quantitatively - that it could not exist anywhere around the atom (as suggested by the Rutherford model), but could instead be found only in specific positions, with specific energies. The electron can move to different positions, but only in discrete, defined steps. It can not rotate anywhere around the nucleus of the atom, but instead was limited to certain areas of space - just as the planets in our solar system are confined in certain ways. Being negatively charged, electrons are attracted to positive protons in the nucleus of the atom and usually occupy an orbital, or path, in the atom that is closest to the nucleus if it is available. This state, which has low energy potential, is called the terrestrial state. By exposing electrons to an external energy source, such as an electrical discharge, electrons can be moved from their state of the earth to other positions that have higher potential energies called excited states. These excited electrons quickly return to lower energy positions (to restore stability associated with lower energies), and energy is released at certain frequencies corresponding to energy differences between electron orbits or shells (see quantum behavior simulation). Bora's mathematical equations further predicted that electrons would not crash into the nucleus in a way that classical physics - and Rutherford's model - predicted. This was another important realization that made the leap from one paradigm (classical physics) to a new (quantum physics). Interactive Animation: The atomic and ionic structure of the first 12 elements, Bohr's discovery that Planck's quantum theory can be applied to the classic Rutherford atom model and may explain the observed flaws in the original model, is another excellent example of how scientific theory uses previous evidence, combined with new experimental observations, to adapt, develop and modify patterns and understand over time. Science is usually advanced by modern scientists to draw on the work of predecessors and, as Isaac Newton put it in his 1676 letter to Robert Hook (both prominent scholars of his time), their standing on the shoulders of giants. Bohr's work is based on the theories of those who were before him, and expanded them to explain the experimentally observed spectra of the lines of atoms in mathematical evidence, which made sense. Showing a checkpoint in the atom, the state of the earth orbital with potential energy. Energy. Bora's work seemed to explain the curious phenomenon of linear spectrums, a number of lines were seen in the hydrogen spectra that did not correspond to Bora's theory. At first glance it seemed to poke holes in Bohr's ideas, but Bohr was quick to offer an explanation. He suggested that the lines in the hydrogen spectrum that cannot be accounted for are actually caused not by hydrogen atoms, but by a completely different species in general. So, what were these different species, and how did they become? Nearly 30 years before Bora published his famous trilogy of articles in a philosophical journal and journal in 1913, the idea of particles having a positive charge, while others have a negative charge. In Bora's work, the ionization of one particular element, helium, proved to be the key to uncovering the mystery of the hydrogen spectrum. When the helium atom, which has two electrons, loses an electron to form helium ion, He+, its electronic structure mimics the atomic hydrogen, since both species have only one electron. Bohr's idea is said to be isoelectronic with a hydrogen atom. However, helium ion has a nucleus with a dual charge of a hydrogen atom and one proton. Bohr understood this and suggested that the greatest part of the electron's mass is located in the nucleus and it falls on the spectral lines that were previously unexplained - the charge of the nucleus affected the energy associated with the transmission of electrons between orbits. Bora's theory was proven correctly when the spectra were created using ionized helium, which was cleaned of hydrogen. At the end of the 19th century, two different particles are known to exist in the atom, and both had an electric charge - a very small and negatively charged electron and a much larger and positively charged proton. By the beginning of the 20th century, however, evidence began to emerge that this was not the full picture of the atom. In particular, the mass of protons and electrons in the atom does not appear to be sufficient to justify the mass of the entire atom, and some types of nuclear decay suggest that something else may occur in the nucleus. 1932. James Chadwick, a British physicist who studied with, and worked for, Ernest Rutherford at the time, set out to solve the problem. Rutherford proposed the idea of a neutral atomic particle, which had mass back in 1920, but he never managed to gain momentum in the hunt for this mysterious particle. In 1932, Chadwick developed an experiment that was first conducted by Frederick Joliot-Curie and Irene Jolio-Curie. They found that by using polonium as a source of alpha particles, they could cause beryllium to emit radiation, which in turn could be used to dislodge protons from a piece of paraffin wax. Joliot-Curie suggested that this radiation was a gamma radiation, a packet of energy without a true mass. As an experienced researcher of gamma rays and the nucleus of the atom, Chadwick realized what others do not have - that the protons are too massive to be ejected from the paraffin of less massive gamma rays. By more carefully measuring the effect of the mysterious particle on paraffin wax and combining this with other measurements, Chadwick concluded that the particles emitted were not gamma radiation, but a relatively heavy particle that had no charge, a particle called a neutron (Figure 5). Figure 5: An artistic model of an atom showing the nucleus, with protons and neutrons, and orbital electrons. Image © Visionlearning Chadwick wrote an article about his discovery entitled Possible Neutron Existence, and it was published in the journal Nature (Chadwick, 1932). In 1935 he was awarded the Nobel Prize in Physics for his discovery. Joliot-Curie also did not go without recognition. Their work on radioactivity and radioactive isotopes won the Nobel Prize in Chemistry, also in 1935. Chadwick's discovery marked the genesis of induced nuclear reactions, where the neutron accelerates and crashes into the nuclei of other elements, generating a huge amount of energy (the atomic bomb). Periodic table elements. German mathematician and physicist Max Bourne made a very specific and critical contribution to quantum mechanics related to the Schroedinger equation. The born took the wave functions that Schroedinger produced, and said that the solutions to the equation could be interpreted as three-dimensional probabilistic maps of where the electron is likely to be found around the atom (Born, 1926). These maps have become known as s, p, d and f orbitals (Figure 3). Figure 3: Based on Bourne's theories, this is a representation of the three-dimensional probabilities of the electron's location around the atom. The four orbits, in increasing complexity, are: s, p, d, and f. Additional information is given about the magnetic quantum number of the orbit (m). Image © Davies/Chemwiki A year after Schroedinger's work was published by German physicist Werner Heisenberg published an article outlining his principle of uncertainty (Heisenberg, 1927). He understands that there are limits to the extent to which the pulse of an electron and its position can be described. Heisenberg's uncertainty principle limits the accuracy of simultaneous knowledge of the particle position and impulse: as the certainty of one increases, then the uncertainty of the other also increases. The most important uncertainty in principle is that with a quantum-mechanical model in which electrons are not in very specific, planetary orbits - the original Bora model - and it is also consistent with Bourne probability maps. Two contributions (Born and Heisenberg) together with the solution of Schroedinger's equation show that the position of the electron in the atom can only be accurately predicted in a statistical way. That is, we know where the electron is most likely found in the atom, but we can never be absolutely sure of its exact position. Understanding of the checkpoint The Heisenberg Uncertainty Principle regarding the position and pulse of the particle states that as the certainty of one increases, so does the number of others. In 1922, German physicist Otto Stern, Bourne's assistant, and Walter Gerlach conducted an experiment in which they passed silver atoms through a magnetic field and observed the deviation model. Simply put, the results gave two different possibilities associated with one, 5s valence electron in each atom. This was an unexpected observation, and implied that one electron could take on two, very different states. At the time, no one could explain the phenomena of the experiment, and it took several scientists working both independently and in unison with earlier experimental observations to work it out for several years. In the early 1920s, the Bohr quantum model and the various spectra that were produced could be adequately described with only three quantum numbers. However, there were experimental observations that could not be explained by only three mathematical parameters. In particular, back in 1896, the Dutch physicist Peter Seaman noted that one valence electron present in the sodium atom can produce two different spectral lines in the presence of a magnetic field. The same phenomenon was observed with other atoms with an odd number of valence electrons. These observations were problematic because they did not fit the working model. In 1925, Dutch physicist George Ulenbeck and his PhD student Samuel Goudsmit suggested that these strange observations could explain whether electrons had an angular impulse; a concept that Wolfgang Pauli later called a spin. As a result, the existence of the fourth quantum number, independent of the orbit on which the electron is located, but unique to a single electron, was revealed. Considering the spin, the idea leads to the next, each of which adds to the refinement of the thinking and understanding, and the advancement of science in this area. The 20th century was a period rich in the development of our knowledge in the field of quantum mechanics, the formation of modern physics. Tracking events during this time, this module covers ideas and refinements that were built on Bohr's groundbreaking work in quantum theory. The contribution of many scientists underscores how theoretical ideas and experimental results have revolutionized our understanding of the atom. Concepts include the Schroedinger equation, Bourne's three-dimensional probability maps, Heisenberg's uncertainty principle and electron spin. The key concepts of Electrons, like light, were shown as wave-like particles, demonstrating the behavior of both waves and particles. Schroedinger's equation describes how the wave particle's wave function changes over time, just as Newton's second law describes the movement of a classical particle. Using quantum numbers, you can write a wave function and find a solution to the equation that helps determine the most likely position of the electron inside the atom. Max Bourne's interpretation of the Schroedinger equation allows us to make three-dimensional probability maps of where electrons can be found around the atom. These maps have become known as s, p, d, and f orbital. Heisenberg's uncertainty principle establishes that the position and pulse of an electron cannot be accurately known together, instead we can only calculate the statistical probability of the electron's location. The opening of the electron rotation determines the fourth quantum number, independent of the electron's orbit, but unique to the electron. The principle of Pauli's exclusion states that no two electrons with the same spin can occupy the same orbital. HS-C1.4, HS-C4.4, HS-PS1. A.2.



We only know about his experiments because they were mentioned in the work of another French chemist and balloonist, Joseph-Louis Guy-Lusac. In 1802, Guy-Lusac published the results of similar experiments, comparing nine different gases. Like Charles, Guy-Lusac came to the conclusion that it is a common property of all gases to increase their volume in the same amount when their temperature has been increased by the same degree. Gay-Lusac kindly gave Charles credit for first observing this general gas behavior. This relationship between gas volume (V) and absolute temperature (T, in Kelvin; to learn more about absolute temperature, see our temperature module) is now known as Charles's Law. Charles's law states that when the pressure remains constant, a fixed amount of gas linearly increases its volume as the temperature rises (Figure 5): Equation 3a Figure 5: Charles's Law states that when the pressure remains constant, a fixed amount of gas linearly increases its volume as the temperature rises. Charles's law can also be understood as: Equation 3b For a fixed amount of gas at a fixed pressure, this constant will be the same even if the volume of gas and temperature vary from (V1, T1) to (V2, T2). Thus, the V1/T1 should be equal to the constant, and the V2/T2 should also be equal to the constant. As a result, the temperature and volume of gas in different conditions are related: Equation 4 This means that if we took a snoopy balloon to the North Pole. The balloon will shrink as the helium cools and decreases in volume. However, if we took the balloon to the hot tropical island and the temperature of helium increased, increased, helium will increase in volume, expanding the balloon. Understanding the checkpoint When different gases are heated by the same number of degrees, their volume will be After his work on Charles's law, Gay-Lussac focused on figuring out how the gases reacted and combined. In 1808, he noticed that many gases combine their volumes into simple, solid ratios. Although we now understand that the volumes of gases are combined in the ratio of all numbers, because this is how gas molecules react, Guy-Lusac did not offer this explanation. This is probably because the idea of molecular combinations of the entire number was only recently proposed by John Dalton, who was Guy-Lusac's scientific rival. (For further study of how gas molecules react, see our chemical equation module). It was the Italian mathematician Amedeo Avogadro who realized that Dalton and Guy-Lusac's ideas complemented each other. Gay-Lusac's assertion that the volumes of gas combined in the ratio of all numbers resembled Dalton's assertion that atoms are combined in the ratio of all numbers to form molecules. Avogadro reasoned that the volume of gas should be related to the number of its molecules. In 1811, Avogadro published his hypothesis that equal volumes of different gases have the same number of molecules. The Avogadro hypothesis was profound, though largely overlooked. The mathematician rarely communicated with other scientists, and he published his hypothesis with mathematical expressions that were unfamiliar to chemists. Nor did it publish experimental data to support his hypothesis. It was 47 years before the Avogadro hypothesis would be widely recognized. In 1858, a former pupil of Avogadro, Italian chemist Stanislao Cannizzaro, published an influential paper on atomic theory. This work relied on the Avogadro hypothesis and presented experimental evidence supporting the hypothesis. The law of Avogadro is based on the Avogadro hypothesis. The Law of Avogadro states that at constant pressure and temperature the volume of gas (V) is directly proportional to the number of molecules. We know that a Snoopy balloon filled with helium will float above the parade, while the same balloon filled with air will trudge on the ground. Although helium and air are very different from the other, Avogadro's law means that if we snuff 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. Understanding the checkpoint Under the Avogadro Act, 1 liter of toxic carbon monoxide gas and 1 liter of flammable hydrogen gas both have the same: Because the gases have a common behavior described by the gas we can understand and predict real gases through the concept of ideal gas - a theoretical, idealized gas that always behaves according to the ideal gas equation. The ideal gas equation stems from gas laws. This equation describes the relationship between all the variables considered in the gas laws: pressure (P), volume (V), quantity (n, in moles) and absolute temperature (T, in Kelvin). Along with the gas constant, R, these variables are combined into an ideal gas equation: Equation 6 Using the ideal gas equation we can solve for any of the unknown variables, as long as we know the others. The R value depends on the units used for other variables (table 2). Table 2: R Gas Constant Values for Different Units. Units R's Value kcal K-1 mol-1 1.9872 J K-1 mol-1 8.3145 L atm K-1 mol-1 0.0821 L Torr K-1 mol-1 62.364 Pa m3 K-1 mol-1 8.3145 Perfect gas law suggests, suggests, that ideal gas molecules have no volume and do not experience intermolecular attraction or repulsion. But real gas molecules have a finite volume and often have some (very small) interactions with each other. However, the behavior and condition of real gas can often be predicted from the ideal gas equation, especially at standard temperature and pressure. In most cases, the difference between real gas behavior and ideal gas behavior is so small that we can use the perfect gas equation for real gases. We will be different from the ideal gas at the end of this module. Understanding Checkpoint Ideal gas molecules have an ideal gas law is also useful in situations where the amount of gas is fixed, but its pressure, volume and temperature change. Using the perfect gas law, we can link the value of these three variables under different conditions. To do this, we must first change the ideal gas equation so that the three changing variables are equal to nR: Equation 7a Equation 7b This relationship is called the Combined Gas Act. Since nR is a constant, we can link the initial (P1, V1, T1) and the final conditions (P2, V2, T2) gas like this: Equation 8 P 1 V 1 T 1 and P 2 V 2 T 2 From 1987 to 2012, air bags (Figure 6) saved nearly 30,000 American lives in car accidents. Air bags save lives because when a car stops hard during an accident, the sensor causes a chemical reaction to generate nitrogen gas. Nitrogen gas inflates the air cushion, which, in fact, forms a large cushion between the driver and the steering wheel. The pillow extends the force of the impact of the accident, helping to reduce the severity of the driver's injuries. For the airbag to work, it must inflate full nitrogen incredibly quickly, within 40 milliseconds after the collision. In order for a 60-liter cylindrical The bag worked properly, the nitrogen gas should reach a pressure of 2.37 37 37 3 m. At 25 degrees Celsius, how many moles of nitrogen gas are needed for Air bag? We can understand this using the perfect gas equation. First, we list the values we know and convert them to have the same units as the gas constant, R (0.0821 L-atm/mol-K). T No. 25 + C (25 th 73 K and 298 K Next, we rearrange the perfect gas equation to decide for the number of moles, n: Finally, we decide for the number of moles of nitrogen gas moles to suppress the air bag: n (2.37 atm ) ( 60 l ) ( 0.0821 litres · am m ka · K ) ( 298 K ) Real gases often deflect off ideal gases when their temperature becomes low, especially when it is close to where the gas will pass a change in its form. When the gas temperature drops, its molecules move more slowly. These slower molecules are less able to overcome even the weak intermolecular forces in the gas. This means that when a gas molecule is about to hit the wall of a container, the very small attraction it experiences for nearby gas molecules reduces its impact and the pressure it exerts on the container. Therefore, real gas at low temperature exerts lower pressure in the container (Figure 7), compared to the ideal gas. Figure 7: Real gas at low temperature exerts lower pressure than predicted due to the attraction between gas molecules. Under high pressure, real gas often deviates from ideal gases because real gas molecules have volume and really attract each other. When real gas is under high pressure, its molecules are more forced to less volume. This smaller volume reduces the amount of free space that molecules have to move (Figure 8). The amount of space that gas molecules take compared to the total space in the container - the relative volume of molecules - becomes larger. Figure 8: Under high pressure, real gas has a larger volume than predicted because of the volume of molecules it contains. This means that at high pressure, real gas has a larger volume than ideal gas, because ideal gas molecules do not have volume. In addition, when real gas molecules are overflowing close together, intermolecular forces can have a greater impact on the behavior of molecules. Attractive intermolecular forces attract molecules to each other, which slows down molecules and reduces their impact on the walls of the container. So when it's under high pressure, real gas has a slightly lower pressure than the ideal gas. Between breaths, the air pressure in our lungs is the same as the atmospheric pressure around us. When we inhale and our chest and diaphragm to expand the volume of our lungs, air pressure decreases and external pressure forces air our lungs until the pressure is the same again, thereby filling our lungs with the oxygen we need to survive. In this module, we focused on the common properties of gases and studied how these properties relate to a common set of behaviors called gas laws. We also got an idea of the ideal gas equation, and when this equation can and cannot be used to predict the behavior of real gases. In other modules, we study the properties of solid and liquid states of matter, as well as study molecular explanations of gas behavior using kinetic molecular theory. This module describes the properties of gases and explores how these properties relate to a common set of behaviors called gas laws. With a focus on Boyle's law, Charles's law, and the Avogadro Law, a review of 400 years of research shows the development of our understanding of gas behavior. The module represents the ideal gas equation and explains when this equation can and cannot be used to predict the behavior of real gases. Key concepts Unlike solids or liquids, molecules in gas are very far apart and rarely interact with each other, so gases made from different molecules have similar behaviors. Gas laws describe the relationship between temperature, pressure, volume and gas. These laws were revealed in experiments carried out by several scientists over four centuries. Because gases have a common behavior, the behavior of real gas at this pressure (P), absolute temperature (T), volume (V) and quantity (n, in moles) can often be predicted by the ideal gas equation, PV and nRT, which perfectly describes the behavior of idealized gas. The behavior of real gases deviates from ideal gases at very low temperatures and high pressure. Ashkenazi, G., James, S.G., and Jason, D.H. (2008). Similarity and difference in the behavior of gases: an interactive demonstration. 85(1): 72. Bell, W.L. (1990). Chemistry of the abducts. 67(1): 11. Brush, S. (1999). Gadflies and geniuses in the history of gas theory. 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The chapter of Robert Boyle's 1660 book with the first experiments on the rarified air. In the Journal of Applied Physiology, 98(1): 31-39. ———. (2014). Robert Hook: early breathing physiologist, polymath and mechanical genius. Physiology (Bethesda), 29 (4): 222-233. Whitaker, R.D. (1979). Early development of kinetic theory. Chemical Education Information Journal, 56 (5): 315. Megan Cartwright, P.H.D., Anthony Carpi, Ph.D. Gas Properties VisionLearning Vol. CHE-3 (9), 2016. Top Page 10 Physical States and Properties by Rachel Bernstein, Ph.D., Anthony Carpi, Ph.D. This is a classic rnap: Fill a salt cocktail with sugar and wait for a meal to take an unexpected turn as your dining companions wonder why their chicken is weirdly sweet. Or go in a different direction and you can really upset someone's morning when they take the first salty sip of coffee. These pranks work well because salt and sugar are almost indistinguishable to the naked eye: Both crystalline solids with similar structures. However, they have very different tastes and they behave differently too. For example, you can pass an electric current through salt water and light a light bulb (you may have done this experiment yourself). But you can't do it with sugar. Differences arise from the different properties of the two crystals, including the atoms that make them up, and the actual structure of the crystal itself. In this module, we will be aware of different types of solids and discuss how their structures relate to their behavior. Figure 1: A pile of salt (left) and sugar (right). From ancient Greece to the birth of modern chemistry in the 17th century, people may have been confused about what made salt and sugar so No modern identification tools components in crystals and their structures, the two would look just as similar to them as they do to our naked eye today (see drawings 1 and 2). As scientists began identifying and characterizing elements in the 17th and 18th centuries, they would have been able to determine that salt is made up of sodium and chlorine, while sugar is made up of carbon, hydrogen and oxygen, but they probably still wondered how such combinations of compounds led to such similar-looking crystals. Figure 2: Typical salt close-up (left) and sugar crystals (right). image © salt levin/dotolery Sugar: Lavi Adler It wasn't until the early 1900s that scientists were first able to look inside crystals when the X-ray diffraction technique was developed. The quartz is a crystalline solid containing a high silicate content (SiO2). If we were to study the structure of quartz, we could see that the silicate units are located very precisely (see Figure 4). Glass, on the other hand, is amorphous solid. While its typical sleek, transparent appearance can make it seem like it should have a neat, organized microscopic structure, on the contrary: silicate units are unevenly scattered all over the solid in a completely disordered fashion. Like quartz, glass has a very high silica content (SiO2). (See our Defining minerals and Silicate Minerals modules for more information on silica and quartz.) The fundamental difference between crystalline and amorphous solids is not what they are made of, but how they are made, but rather how their structures are structured. The quartz is formed on a very slow, geological timeline, so that atoms have time to reach a highly ordered crystalline structure in which atoms optimize attractive forces and minimize the repulsive forces between them and which are therefore energetically favorable. Glass, on the other hand, is made by melting sand (among other methods) and letting it cool very quickly, freezing the atoms in place, causing disorderly amorphous solids. Amorphous solids are often when atoms and molecules are frozen in place before they have a chance to reach a crystalline arrangement, which would otherwise be the preferred structure because it is energy-favored. One important consequence of the improper structure of amorphous solids is that they always behave consistently or evenly. For example, they can melt in a wide range of temperatures, as opposed to a very precise melting point for crystalline solids. Returning to the example of glass and quartz, the most common type of glass, called soda lime glass, can melt somewhere between 550 and 1450 degrees Celsius, while the cristobalite, a quartz polymorph, melts exactly at 1713 degrees Celsius. In addition, amorphous solids break unpredictably and produce fragments with irregular, often curved surfaces, while crystalline solids break down along certain planes and at certain angles, defined by the geometry of the crystal. (See our Defining Minerals module for more information on how the appearance of the crystal reflects the regular location of its atoms.) Understanding checkpoint How amorphous solid, glass has an exact melting point. The crystal structure determines much more about solid than just how it breaks down. The structure is directly related to a number of important properties, including, for example, conductivity and density, in particular. To explain these relationships, we first have to introduce four main types of crystalline solids - molecular, network, ion and metal - each of which is described below. Individual molecules consist of atoms held together by strong covalent bonds (see our chemical bonding module for more information on covalent bonding). To form molecular solids, these molecules are then arranged in a regular, repeating pattern. Consider the following examples: (1) Diamond and graphite. Both are made of carbon atoms, but they have different structures. In diamond, each carbon atom is tetrahedrally bonded to four other carbon atoms, forming a three-dimensional network of strong covalent bonds. In graphite, carbon atoms are arranged in layers, with each carbon atom bonded to three other carbon atoms in the same plane, forming a two-dimensional network. The difference in structure leads to very different properties. Diamond is extremely hard and a good conductor of electricity, while graphite is soft and a good conductor of electricity. (2) Silicon and germanium. Both are made of silicon and germanium atoms, but they have different structures. In silicon, each atom is tetrahedrally bonded to four other silicon atoms, forming a three-dimensional network. In germanium, each atom is tetrahedrally bonded to four other germanium atoms, forming a three-dimensional network. The difference in structure leads to very different properties. Silicon is a semiconductor, while germanium is a semiconductor. (3) Sodium chloride and potassium chloride. Both are made of sodium and chlorine atoms, but they have different structures. In sodium chloride, each sodium atom is surrounded by six chlorine atoms, forming a cubic lattice. In potassium chloride, each potassium atom is surrounded by six chlorine atoms, forming a cubic lattice. The difference in structure leads to very different properties. Sodium chloride is a common salt, while potassium chloride is used in medicine. (4) Water and ice. Both are made of hydrogen and oxygen atoms, but they have different structures. In water, each water molecule is surrounded by four other water molecules, forming a tetrahedral network. In ice, each water molecule is surrounded by four other water molecules, forming a tetrahedral network. The difference in structure leads to very different properties. Water is a liquid, while ice is a solid. (5) Metals. Metals are made of metal atoms, but they have different structures. In metals, each atom is surrounded by a sea of delocalized electrons, forming a metallic lattice. The difference in structure leads to very different properties. Metals are good conductors of electricity and heat, and they are malleable and ductile. (6) Molecular solids. Molecular solids are made of molecules, but they have different structures. In molecular solids, each molecule is surrounded by other molecules, forming a regular lattice. The difference in structure leads to very different properties. Molecular solids can be hard or soft, and they can be good conductors or insulators. (7) Network solids. Network solids are made of atoms, but they have different structures. In network solids, each atom is bonded to other atoms, forming a continuous network of covalent bonds. Network solids are hard and have high melting points. (8) Ionic solids. Ionic solids are made of ions, but they have different structures. In ionic solids, each ion is surrounded by other ions, forming a regular lattice. Ionic solids are hard and have high melting points. (9) Amorphous solids. Amorphous solids are made of atoms or molecules, but they have different structures. In amorphous solids, the atoms or molecules are not arranged in a regular lattice. Amorphous solids are soft and have low melting points. (10) Polymers. Polymers are made of long chains of repeating units, but they have different structures. In polymers, the repeating units are connected by covalent bonds, forming a long chain. Polymers can be hard or soft, and they can be good conductors or insulators. (11) Liquid crystals. Liquid crystals are made of molecules, but they have different structures. In liquid crystals, the molecules are arranged in a regular lattice, but they can flow like a liquid. Liquid crystals are used in displays and optical devices. (12) Gels. Gels are made of polymers, but they have different structures. In gels, the polymer chains are cross-linked, forming a network. Gels are used in many applications, including food, medicine, and materials science. (13) Nanomaterials. Nanomaterials are made of atoms or molecules, but they have different structures. In nanomaterials, the atoms or molecules are arranged in a regular lattice, but they have unique properties due to their small size. Nanomaterials are used in many applications, including electronics, medicine, and materials science. (14) Composites. Composites are made of two or more materials, but they have different structures. In composites, the materials are combined to form a new material with unique properties. Composites are used in many applications, including aerospace, construction, and sports equipment. (15) Biomaterials. Biomaterials are made of natural or synthetic materials, but they have different structures. In biomaterials, the materials are used to replace or repair a part of the body. Biomaterials are used in many applications, including implants, prosthetics, and tissue engineering. (16) Smart materials. Smart materials are made of materials that can change their properties in response to an external stimulus. Smart materials are used in many applications, including sensors, actuators, and robotics. (17) Metamaterials. Metamaterials are made of materials that have properties not found in nature. Metamaterials are used in many applications, including隐身斗篷, superlenses, and antennas. (18) Quantum materials. Quantum materials are made of materials that exhibit quantum effects. Quantum materials are used in many applications, including quantum computing, quantum cryptography, and quantum communication. (19) Topological materials. Topological materials are made of materials that have unique topological properties. Topological materials are used in many applications, including quantum computing, quantum cryptography, and quantum communication. (20) 2D materials. 2D materials are made of materials that are only two atoms thick. 2D materials are used in many applications, including electronics, optics, and energy storage. (21) Nanowires. Nanowires are made of materials that are only a few nanometers wide. Nanowires are used in many applications, including electronics, optics, and energy storage. (22) Quantum dots. Quantum dots are made of materials that are only a few nanometers in size. Quantum dots are used in many applications, including electronics, optics, and energy storage. (23) Carbon nanotubes. Carbon nanotubes are made of carbon atoms arranged in a cylindrical lattice. Carbon nanotubes are used in many applications, including electronics, optics, and energy storage. (24) Graphene. Graphene is a single layer of carbon atoms arranged in a hexagonal lattice. Graphene is used in many applications, including electronics, optics, and energy storage. (25) Nanoparticles. Nanoparticles are made of materials that are only a few nanometers in size. Nanoparticles are used in many applications, including medicine, electronics, and energy storage. (26) Nanofibers. Nanofibers are made of materials that are only a few nanometers in diameter. Nanofibers are used in many applications, including medicine, electronics, and energy storage. (27) Nanosheets. 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