Solid

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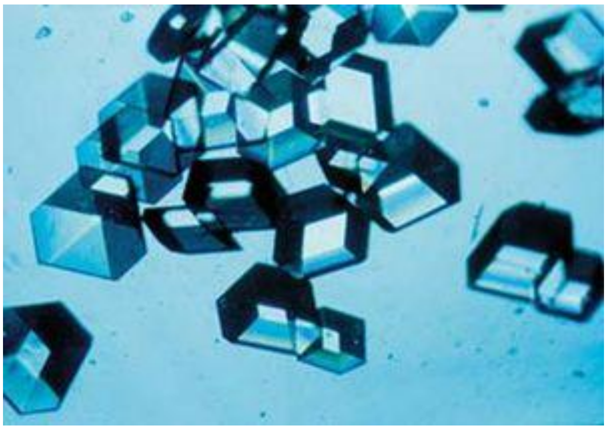
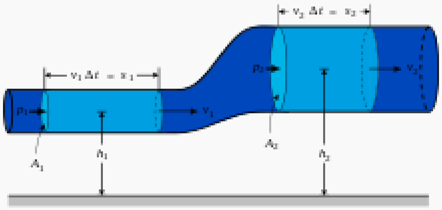
For other uses, see Solid (disambiguation).



Single crystalline form of solid Insulin.

Continuum mechanics



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Solid is one of the four fundamental states of matter (the others being liquid, gas, and plasma). It is characterized by structural rigidity and resistance to changes of shape or volume. Unlike a liquid, a solid object does not flow to take on the shape of its container, nor does it expand to fill the entire volume available to it like a gas does. The atoms in a solid are tightly bound to each other, either in a regular geometric lattice (crystalline solids, which include metals and ordinary water ice) or irregularly (an amorphous solid such as common window glass).

The branch of physics that deals with solids is called solid-state physics, and is the main branch of condensed matter physics (which also includes liquids). Materials science is primarily concerned with the physical and chemical properties of solids. Solid-state chemistry is especially concerned with the synthesis of novel materials, as well as the science of identification and chemical composition.

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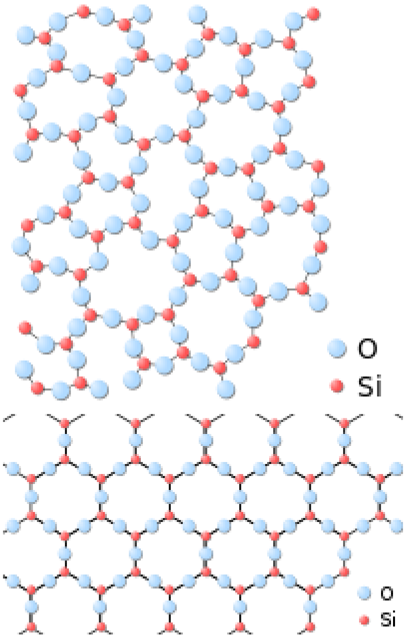
[edit] Microscopic description

Model of closely packed atoms within a crystalline solid.

The atoms, molecules or ions which make up a solid may be arranged in an orderly repeating pattern, or irregularly. Materials whose constituents are arranged in a regular pattern are known as crystals. In some cases, the regular ordering can continue unbroken over a large scale, for example diamonds, where each diamond is a single crystal. Solid objects that are large enough to see and handle are rarely composed of a single crystal, but instead are made of a large number of single crystals, known as crystallites, whose size can vary from a few nanometers to several meters. Such materials are called polycrystalline. Almost all common metals, and many ceramics, are polycrystalline.



Schematic representation of a random-network glassy form (left) and ordered crystalline lattice (right) of identical chemical composition.

In other materials, there is no long-range order in the position of the atoms. These solids are known as amorphous solids; examples include polystyrene and glass.

Whether a solid is crystalline or amorphous depends on the material involved, and the conditions in which it was formed. Solids which are formed by slow cooling will tend to be crystalline, while solids which are frozen rapidly are more likely to be amorphous. Likewise, the specific crystal structure adopted by a crystalline solid depends on the material involved and on how it was formed.

While many common objects, such as an ice cube or a coin, are chemically identical throughout, many other common materials comprise a number of different substances packed together. For example, a typical rock is an aggregate of several different minerals and mineraloids, with no specific chemical composition. Wood is a natural organic material consisting primarily of cellulose fibers embedded in a matrix of organic lignin. In materials science, composites of more than one constituent material can be designed to have desired properties.

[edit] Classes of solids

Further information: Bonding in solids

The forces between the atoms in a solid can take a variety of forms. For example, a crystal of sodium chloride (common salt) is made up of ionic sodium and chlorine, which are held together by ionic bonds. In diamond or silicon, the atoms share electrons and form covalent bonds. In metals, electrons are shared in metallic bonding. Some solids, particularly most organic compounds, are held together with van der Waals forces resulting from the polarization of the electronic charge cloud on each molecule. The dissimilarities between the types of solid result from the differences between their bonding.

[edit] Metals

Main article: Metal

The pinnacle of New York's Chrysler Building, the world's tallest steel-supported brick building, is clad with stainless steel.

Metals typically are strong, dense, and good conductors of both electricity and heat. The bulk of the elements in the periodic table, those to the left of a diagonal line drawn from boron to polonium, are metals. Mixtures of two or more elements in which the major component is a metal are known as alloys.

People have been using metals for a variety of purposes since prehistoric times. The strength and reliability of metals has led to their widespread use in construction of buildings and other structures, as well as in most vehicles, many appliances and tools, pipes, road signs and railroad tracks. Iron and aluminium are the two most commonly used structural metals, and they are also the most abundant metals in the Earth's crust. Iron is most commonly used in the form of an alloy, steel, which contains up to 2.1% carbon, making it much harder than pure iron.

Because metals are good conductors of electricity, they are valuable in electrical appliances and for carrying an electric current over long distances with little energy loss or dissipation. Thus, electrical power grids rely on metal cables to distribute electricity. Home electrical systems, for example, are wired with copper for its good conducting properties and easy machinability. The high thermal conductivity of most metals also makes them useful for stovetop cooking utensils.

The study of metallic elements and their alloys makes up a significant portion of the fields of solid-state chemistry, physics, materials science and engineering.

Metallic solids are held together by a high density of shared, delocalized electrons, known as "metallic bonding". In a metal, atoms readily lose their outermost ("valence") electrons, forming positive ions. The free electrons are spread over the entire solid, which is held together firmly by electrostatic interactions between the ions and the electron cloud.[1] The large number of free electrons gives metals their high values of electrical and thermal conductivity. The free electrons also prevent transmission of visible light, making metals opaque, shiny and lustrous.

More advanced models of metal properties consider the effect of the positive ions cores on the delocalised electrons. As most metals have crystalline structure, those ions are usually arranged into a periodic lattice. Mathematically, the potential of the ion cores can be treated by various models, the simplest being the nearly free electron model.

[edit] Minerals

A collection of various minerals.

Main article: Minerals

Minerals are naturally occurring solids formed through various geological processes under high pressures. To be classified as a true mineral, a substance must have a crystal structure with uniform physical properties throughout. Minerals range in composition from pure elements and simple salts to very complex silicates with thousands of known forms. In contrast, a rock sample is a random aggregate of minerals and/or mineraloids, and has no specific chemical composition. The vast majority of the rocks of the Earth's crust consist of quartz (crystalline SiO_2), feldspar, mica, chlorite, kaolin, calcite, epidote, olivine, augite, hornblende, magnetite, hematite, limonite and a few other minerals. Some minerals, like quartz, mica or feldspar are common, while others have been found in only a few locations worldwide. The largest group of minerals by far is the silicates (most rocks are ≥95% silicates), which are composed largely of silicon and oxygen, with the addition of ions of aluminium, magnesium, iron, calcium and other metals.

[edit] Ceramics

Si_3N_4 ceramic bearing parts

Main article: Ceramic engineering

Ceramic solids are composed of inorganic compounds, usually oxides of chemical elements. They are chemically inert, and often are capable of withstanding chemical erosion that occurs in an acidic or caustic environment. Ceramics generally can withstand high temperatures ranging from 1000 to 1600 °C (1800 to 3000 °F). Exceptions include non-oxide inorganic materials, such as nitrides, borides and carbides.

Traditional ceramic raw materials include clay minerals such as kaolinite, more recent materials include aluminium oxide (alumina). The modern ceramic

Liquid

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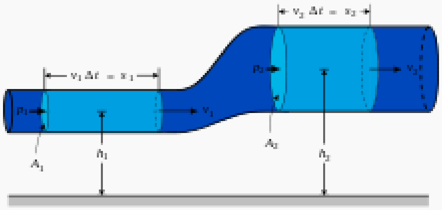
For other uses, see Liquid (disambiguation).

The formation of a spherical droplet of liquid water minimizes the surface area, which is the natural result of surface tension in liquids.

Continuum mechanics



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Liquid is one of the the four fundamental states of matter (the others being solid, gas, and plasma), and is the only state with a definite volume but no fixed shape. A liquid is made up of tiny vibrating particles of matter, such as atoms and molecules, held together by intramolecular bonds. Water is, by far, the most common liquid on Earth. Like a gas, a liquid is able to flow and take the shape of a container. Some liquids resist compression, while others can be compressed. Unlike a gas, a liquid does not disperse to fill every space of a container, and maintains a fairly constant density. A distinctive property of the liquid state is surface tension, leading to wetting phenomena.

The density of a liquid is usually close to that of a solid, and much higher than in a gas. Therefore, liquid and solid are both termed condensed matter. On the other hand, as liquids and gases share the ability to flow, they are both called fluids. Although liquid water is abundant on Earth, this state of matter is actually the least common in the known universe, because liquids require a relatively narrow temperature/pressure range to exist. Most known matter in the universe is in gaseous form (with traces of detectable solid matter) as interstellar clouds or in plasma form within stars.

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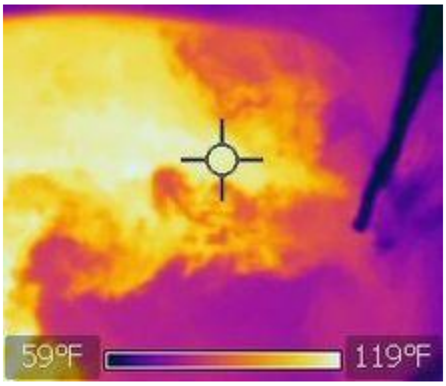
[edit] Introduction



Thermal image of a sink full of hot water with cold water being added, showing how the hot and the cold water flow into each other.

Liquid is one of the three primary states of matter, with the others being solid and gas. A liquid is a fluid. Unlike a solid, the molecules in a liquid have a much greater freedom to move. The forces that bind the molecules together in a solid are only temporary in a liquid, allowing a liquid to flow while a solid remains rigid.

A liquid, like a gas, displays the properties of a fluid. A liquid can flow, assume the shape of a container, and, if placed in a sealed container, will distribute applied pressure evenly to every surface in the container. Unlike a gas, a liquid may not always mix readily with another liquid, will not always fill every space in the container, forming its own surface, and will not compress significantly, except under extremely high pressures. These properties make a liquid suitable for applications such as hydraulics.

Liquid particles are bound firmly but not rigidly. They are able to move around one another freely, resulting in a limited degree of particle mobility. As the temperature increases, the increased vibrations of the molecules causes distances between the molecules to increase. When a liquid reaches its boiling point, the cohesive forces that bind the molecules closely together break, and the liquid changes to its gaseous state (unless superheating occurs). If the temperature is decreased, the distances between the molecules become smaller. When the liquid reaches its freezing point the molecules will usually lock into a very specific order, called crystallizing, and the bonds between them become more rigid, changing the liquid into its solid state (unless supercooling occurs).

[edit] Examples

Only two elements are liquid at standard conditions for temperature and pressure: mercury and bromine. Four more elements have melting points slightly above room temperature: francium, caesium, gallium and rubidium.[1] Metal alloys that are liquid at room temperature include NaK, a sodium-potassium metal alloy, galinstan, a fusible alloy liquid, and some amalgams (alloys involving mercury).

Pure substances that are liquid under normal conditions include water, ethanol and many other organic solvents. Liquid water is of vital importance in chemistry and biology; it is believed to be a necessity for the existence of life.

Important everyday liquids include aqueous solutions like household bleach, other mixtures of different substances such as mineral oil and gasoline, emulsions like vinaigrette or mayonnaise, suspensions like blood, and colloids like paint and milk.

Many gases can be liquefied by cooling, producing liquids such as liquid oxygen, liquid nitrogen, liquid hydrogen and liquid helium. Not all gases can be liquified at atmospheric pressure, for example carbon dioxide can only be liquified at pressures above 5.1 atm.

Some materials cannot be classified within the classical three states of matter; they possess solid-like and liquid-like properties. Examples include liquid crystals, used in LCD displays, and biological membranes.

[edit] Applications

Liquids have a variety of uses, as lubricants, solvents, and coolants. In hydraulic systems, liquid is used to transmit power.

In tribology, liquids are studied for their properties as lubricants. Lubricants such as oil are chosen for viscosity and flow characteristics that are suitable throughout the operating temperature range of the component. Oils are often used in engines, gear boxes, metalworking, and hydraulic systems for their good lubrication properties.[2]

Many liquids are used as solvents, to dissolve other liquids or solids. Solutions are found in a wide variety of applications, including paints, sealants, and adhesives. Naptha and acetone are used frequently in industry to clean oil, grease, and tar from parts and machinery. Body fluids are water based solutions.

Surfactants are commonly found in soaps and detergents. Solvents like alcohol are often used as antimicrobials. They are found in cosmetics, inks, and liquid dye lasers. They are used in the food industry, in processes such as the extraction of vegetable oil.[3]

Liquids tend to have better thermal conductivity than gases, and the ability to flow makes a liquid suitable for removing excess heat from mechanical components. The heat can be removed by channeling the liquid through a heat exchanger, such as a radiator, or the heat can be removed with the liquid during evaporation.[4] Water or glycol coolants are used to keep engines from overheating.[5] The coolants used in nuclear reactors include water or liquid metals, such as sodium or bismuth.[6] Liquid propellant films are used to cool the thrust chambers of rockets.[7] In machining, water and oils are used to remove the excess heat generated, which can quickly ruin both the work piece and the tooling. During perspiration, sweat removes heat from the human body by evaporating. In the heating, ventilation, and air-conditioning industry (HVAC), liquids such as water are used to transfer heat from one area to another.[8]

Liquid is the primary component of hydraulic systems, which take advantage of Pascal's law to provide fluid power. Devices such as pumps and waterwheels have been used to change liquid motion into mechanical work since ancient times. Oils are forced through hydraulic pumps, which transmit this force to hydraulic cylinders. Hydraulics can be found in many applications, such as automotive brakes and transmissions, heavy equipment, and airplane control systems. Various hydraulic presses are used extensively in repair and manufacturing, for lifting, pressing, clamping and forming.[9]

Liquids are sometimes used in measuring devices. A thermometer often uses the thermal expansion of liquids, such as mercury, combined with their ability to flow to indicate temperature. A manometer uses the weight of the liquid to indicate air pressure.[10]

[edit] Mechanical properties

[edit] Volume

Quantities of liquids are commonly measured in units of volume. These include the SI unit cubic metre (m^3) and its divisions, in particular the cubic decimetre, more commonly called the litre ($1\ dm^3 = 1\ L = 0.001\ m^3$), and the cubic centimetre, also called millilitre ($1\ cm^3 = 1\ mL = 0.001\ L = 10^{-6}\ m^3$).

The volume of a quantity of liquid is fixed by its temperature and pressure. Liquids generally expand when heated, and contract when cooled. Water between 0 °C and 4 °C is a notable exception. Liquids have little compressibility: water, for example, requires a pressure of the order of 200 bar to increase its density by 1/1000. In the study of fluid dynamics, liquids are often treated as incompressible, especially when studying incompressible flow.

[edit] Pressure and buoyancy

Main article: fluid statics

In a gravitational field, liquids exert pressure on the sides of a container as well as on anything within the liquid itself. This pressure is transmitted in all directions and increases with depth. If a liquid is at rest in a uniform gravitational field, the pressure, p, at any depth, z, is given by

$$p = \rho g z$$

where:

ρ is the density of the liquid (assumed constant)
g is the gravitational acceleration.

Note that this formula assumes that the pressure *at* the free surface is zero, and that surface tension effects may be neglected.

Objects immersed in liquids are subject to the phenomenon of buoyancy. (Buoyancy is also observed in other fluids, but is especially strong in liquids due to their high density.)

[edit] Surfaces

Main article: surface science

Surface waves in water

Unless the volume of a liquid exactly matches the volume of its container, one or more surfaces are observed. The surface of a liquid behaves like an elastic membrane in which surface tension appears, allowing the formation of drops and bubbles. Surface waves, capillary action, wetting, and ripples are other consequences of surface tension.

[edit] Flow

Main article: fluid mechanics

Main article: fluid dynamics

Viscosity measures the resistance of a liquid which is being deformed by either shear stress or extensional stress.

When a liquid is supercooled towards the glass transition, the viscosity increases dramatically! The liquid then becomes a viscoelastic medium that shows both the elasticity of a solid and the fluidity of a liquid, depending on the time scale of observation or on the frequency of perturbation.

[edit] Sound propagation

Main article: speed of sound#Speed of sound in liquids

In a fluid the only non-zero stiffness is to volumetric deformation (a fluid does not sustain shear forces). Hence the speed of sound in a fluid is given by $c = \sqrt{K/\rho}$ where K is the bulk modulus of the fluid, and ρ the density. To give a typical value, in fresh water c=1497 m/s at 25 °C.

[edit] Thermodynamics

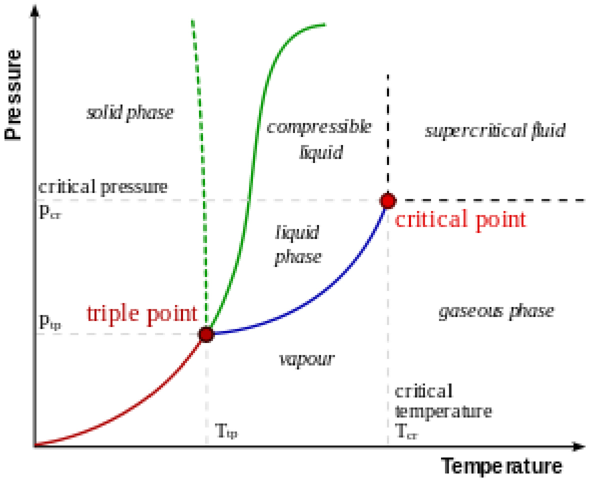
[edit] Phase transitions

Main article: boiling

Main article: boiling point

Main article: melting

Main article: melting point



A typical phase diagram. The dotted line gives the anomalous behaviour of water. The green lines show how the freezing point can vary with pressure, and the blue line shows how the boiling point can vary with pressure. The red line shows the boundary where sublimation or deposition can occur.

At a temperature below the boiling point, any matter in liquid form will evaporate until the condensation of gas above reach an equilibrium. At this point the gas will condense at the same rate as the liquid evaporates. Thus, a liquid cannot exist permanently if the evaporated liquid is continually removed. A liquid at its boiling point will evaporate more quickly than the gas can condense at the current pressure. A liquid at or above its boiling point will normally boil, though superheating can prevent this in certain circumstances.

At a temperature below the freezing point, a liquid will tend to crystallize, changing to its solid form. Unlike the transition to gas, there is no equilibrium at this transition under constant pressure, so unless supercooling occurs, the liquid will eventually completely crystallize. Note that this is only true under constant pressure, so e.g. water and ice in a closed, strong container might reach an equilibrium where both phases coexist. For the opposite transition from solid to liquid, see melting.

[edit] Solutions

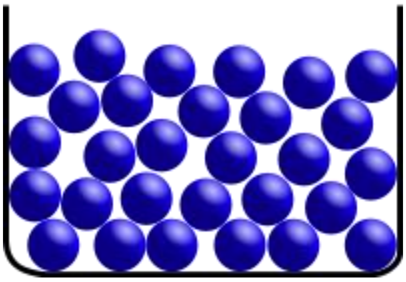
Main article: solution

Liquids can display immiscibility. The most familiar mixture of two immiscible liquids in everyday life is the vegetable oil and water in Italian salad dressing. A familiar set of miscible liquids is water and alcohol. Liquid components in a mixture can often be separated from one another via fractional distillation.

[edit] Microscopic properties

[edit] Static structure factor

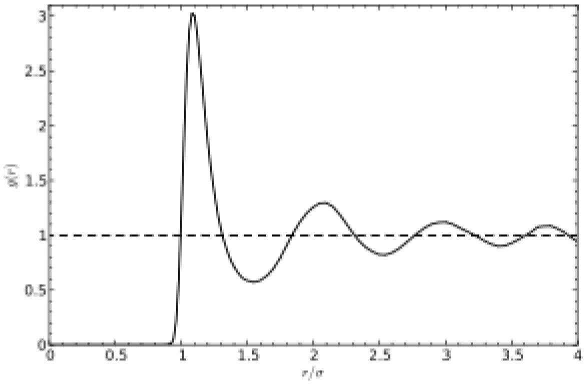
Main article: structure of liquids and glasses



Structure of a classical monatomic liquid. Atoms have many nearest neighbors in contact, yet no long-range order is present.

In a liquid, atoms do not form a crystalline lattice, nor do they show any other form of long-range order. This is evidenced by the absence of Bragg peaks in X-ray and neutron diffraction. Under normal conditions, the diffraction pattern has circular symmetry, expressing the isotropy of the liquid. In radial direction, the diffraction intensity smoothly oscillates. This is usually described by the static structure factor *S(q)*, with wavenumber *q*=(4π/λ)sinθ given by the wavelength λ of the probe (photon or neutron) and the Bragg angle θ. The oscillations of *S(q)* express the *near order* of the liquid, i.e. the correlations between an atom and a few shells of nearest, second nearest, ... neighbors.

A more intuitive description of these correlations is given by the radial distribution function *g(r)*, which is basically the Fourier transform of *S(q)*. It represents a spatial average of a temporal snapshot of pair correlations in the liquid.



Radial distribution function of the Lennard-Jones model fluid.

[edit] Sound dispersion and structural relaxation

The above expression for the sound velocity $c = \sqrt{K/\rho}$contains the bulk modulus *K*. If *K* is frequency independent then the liquid behaves as a linear medium, so that sound propagates without dissipation and without mode coupling. In reality, any liquid shows some dispersion: with increasing frequency, *K* crosses over from the low-frequency, liquid-like limit $K_0$to the high-frequency, solid-like limit K_∞. In normal liquids, most of this cross over takes place at frequencies between GHz and THz, sometimes called hypersound.

At sub-GHz frequencies, a normal liquid cannot sustain shear waves: the zero-frequency limit of the shear modulus is $G_0 = 0$. This is sometimes seen as the defining property of a liquid.[11] However, just as the bulk modulus K, the shear modulus G is frequency dependent, and at hypersound frequencies it shows a similar cross over from the liquid-like limit G_0 to a solid-like, non-zero limit G_∞.

According to the Kramers-Kronig relation, the dispersion in the sound velocity (given by the real part of K or G) goes along with a maximum in the sound attenuation (dissipation, given by the imaginary part of K or G). According to linear response theory, the Fourier transform of K or G describes how the system returns to equilibrium after an external perturbation; for this reason, the dispersion step in the GHz..THz region is also called structural relaxation. According the fluctuation-dissipation theorem, relaxation *towards* equilibrium is intimately connected to fluctuations *in* equilibrium. The density fluctuations associated with sound waves can be experimentally observed by Brillouin scattering.

On supercooling a liquid towards the glass transition, the crossover from liquid-like to solid-like response moves from GHz to MHz, kHz, Hz, ...; equivalently, the characteristic time of structural relaxation increases from ns to µs, ms, s, ... This is the microscopic explanation for the above mentioned viscoelastic behaviour of glass-forming liquids.

[edit] Effects of association

The mechanisms of atomic/molecular diffusion (or particle displacement) in solids are closely related to the mechanisms of viscous flow and solidification in liquid materials. Descriptions of viscosity in terms of molecular "free space" within the liquid[12] were modified as needed in order to account for liquids whose molecules are known to be "associated" in the liquid state at ordinary temperatures. When various molecules combine together to form an associated molecule, they enclose within a semi-rigid system a certain amount of space which before was available as free space for mobile molecules. Thus, increase in viscosity upon cooling due to the tendency of most substances to become *associated* on cooling.[13]

Similar arguments could be used to describe the effects of pressure on viscosity, where it may be assumed that the viscosity is chiefly a function of the volume for liquids with a finite compressibility. An increasing viscosity with rise of pressure is therefore expected. In addition, if the volume is expanded by heat but reduced again by pressure, the viscosity remains the same.

The local tendency to orientation of molecules in small groups lends the liquid (as referred to previously) a certain degree of association. This association results in a considerable "internal pressure" within a liquid, which is due almost entirely to those molecules which, on account of their temporary low velocities (following the Maxwell distribution) have coalesced with other molecules. The internal pressure between several such molecules might correspond to that between a group of molecules in the solid form.

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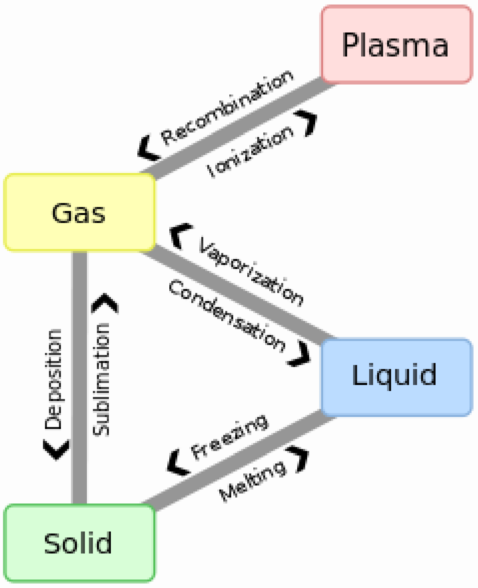
States of matter (classic)

State	Solid · **Liquid** · Gas / Vapor · Plasma
Low energy	Bose–Einstein condensate · Fermionic condensate · Degenerate matter · Quantum Hall · Rydberg matter · Strange matter ·



	Superfluid · Supersolid
High energy	QCD matter · Quark–gluon plasma · Supercritical fluid
Other states	Colloid · Glass · Liquid crystal · Magnetically ordered (Antiferromagnet · Ferrimagnet · Ferromagnet) · String-net liquid · Superglass
Transitions	Boiling · Boiling point · Condensation · Critical line · Critical point · Crystallization · Deposition · Evaporation ·

	- Flash evaporation · - Freezing · - Ionization · - Lambda point · - Melting · - Melting point · - Recombination · - Regelation · - Saturated fluid · - Sublimation · - Supercooling · - Triple point · - Vaporization · - Vitrification
Quantities	- Enthalpy of fusion · - Enthalpy of sublimation · - Enthalpy of vaporization · - Latent heat · - Latent internal energy · - Trouton's ratio · - Volatility

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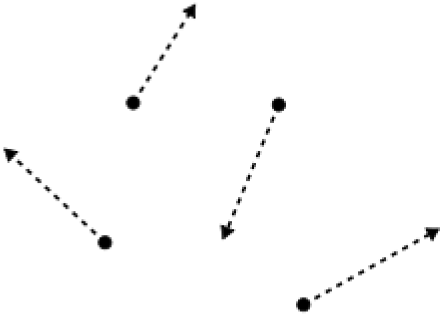
Gas

From Wikipedia, the free encyclopedia

Jump to: navigation, search

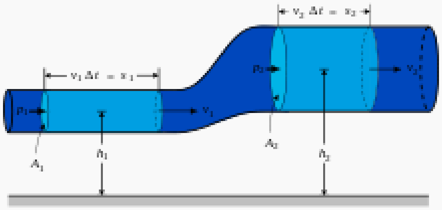
This article is about the physical properties of gas as a state of matter. For the uses of gases, and other meanings, see Gas (disambiguation).

This article **is written like a manual or guidebook.** Please help rewrite this article from a descriptive, neutral point of view, and remove advice or instruction. If it is intended that a manual be produced, use of either WikiHow or Wikibooks is strongly suggested. *(January 2010)*



Gas phase particles (atoms, molecules, or ions) move around freely in the absence of an applied electric field.

Continuum mechanics



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Gas is one of the four fundamental states of matter (the others being solid, liquid, and plasma). A pure gas may be made up of individual atoms (e.g. a noble gas or atomic gas like neon), elemental molecules made from one type of atom (e.g. oxygen), or compound molecules made from a variety of atoms (e.g. carbon dioxide). A gas mixture would contain a variety of pure gases much like the air. What distinguishes a gas from liquids and solids is the vast separation of the individual gas particles. This separation usually makes a colorless gas invisible to the human observer. The interaction of gas particles in the presence of electric and gravitational fields are considered negligible as indicated by the constant velocity vectors in the image.

The gaseous state of matter is found between the liquid and plasma states,[1] the latter of which provides the upper temperature boundary for gases. Bounding the lower end of the temperature scale lie degenerative quantum gases[2] which are gaining increased attention these days.[3] High-density atomic gases super cooled to incredibly low temperatures are classified by their statistical behavior as either a Bose gas or a Fermi gas. For a comprehensive listing of these exotic states of matter see list of states of matter.

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Etymology

The word *gas* is a neologism first used by the early 17th century Flemish chemist J.B. Van Helmont.[4] Van Helmont's word appears to have been simply a phonetic transcription of the Greek word χάος *Chaos* – the *g* in Dutch being pronounced like the English *ch* – in which case Van Helmont was simply following the established alchemical usage first attested in the works of Paracelsus. According to Paracelsus's terminology, *chaos* meant something like "ultra-rarefied water".[5]

Physical characteristics

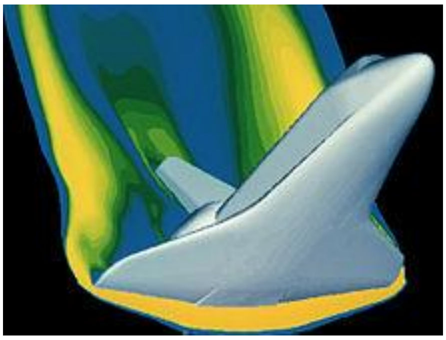
Drifting smoke particles provide clues to the movement of the surrounding gas.

As most gases are difficult to observe directly, they are described through the use of four physical properties or macroscopic characteristics: pressure, volume, number of particles (chemists group them by moles) and temperature. These four characteristics were repeatedly observed by scientists such as Robert Boyle, Jacques Charles, John Dalton, Joseph Gay-Lussac and Amedeo Avogadro for a variety of gases in various settings. Their detailed studies ultimately led to a mathematical relationship among these properties expressed by the ideal gas law (see simplified models section below).

Gas particles are widely separated from one another, and consequently have weaker intermolecular bonds than liquids or solids. These intermolecular forces result from electrostatic interactions between gas particles. Like-charged areas of different gas particles repel, while oppositely charged regions of different gas particles attract one another; gases that contain permanently charged ions are known as plasmas. Gaseous compounds with polar covalent bonds contain permanent charge imbalances and so experience relatively strong intermolecular forces, although the molecule while the compound's net charge remains neutral. Transient, randomly-induced charges exist across non-polar covalent bonds of molecules and electrostatic interactions caused by them are referred to as Van der Waals forces. The interaction of these intermolecular forces varies within a substance which determines many of the physical properties unique to each gas.[6][7] A comparison of *boiling points* for compounds formed by ionic and covalent bonds leads us to this conclusion.[8] The drifting smoke particles in the image provides some insight into low pressure gas behavior.

Compared to the other states of matter, gases have low density and viscosity. Pressure and temperature influence the particles within a certain volume. This variation in particle separation and speed is referred to as *compressibility*. This particle separation and size influences optical properties of gases as can be found in the following list of refractive indices. Finally, gas particles spread apart or diffuse in order to homogeneously distribute themselves throughout any container.

Macroscopic



Shuttle imagery of re-entry phase.

When observing a gas, it is typical to specify a frame of reference or length scale. A *larger* length scale corresponds to a macroscopic or global point of view of the gas. This region (referred to as a volume) must be sufficient in size to contain a large sampling of gas particles. The resulting statistical analysis of this sample size produces the **"average"** behavior (i.e. velocity, temperature or pressure) of all the gas particles within the region. In contrast, a *smaller* length scale corresponds to a microscopic or particle point of view.

Macroscopically, the gas characteristics measured are either in terms of the gas particles themselves (velocity, pressure, or temperature) or their surroundings (volume). For example, Robert Boyle studied pneumatic chemistry for a small portion of his career. One of his experiments related the macroscopic properties of pressure and volume of a gas. His experiment used a J-tube manometer which looks like a test tube in the shape of the letter J. Boyle trapped an inert gas in the closed end of the test tube with a column of mercury, thereby making the number of particles and the temperature constant. He observed that when the pressure was increased in the gas, by adding more mercury to the column, the trapped gas' volume decreased (this is known as an inverse relationship). Furthermore, when Boyle multiplied the pressure and volume of each observation, the product was constant. This relationship held for every gas that Boyle observed leading to the law, (PV=k), named to honor his work in this field.

There are many mathematical tools available for analyzing gas properties. As gases are subjected to extreme conditions, these tools become a bit more complex, from the Euler equations for inviscid flow to the Navier-Stokes equations[9] that fully account for viscous effects. These equations are adapted to the conditions of the gas system in question. Boyle's lab equipment allowed the use of algebra to obtain his analytical results. His results were possible because he was studying gases in relatively low pressure situations where they behaved in an "ideal" manner. These ideal relationships apply to safety calculations for a variety of flight conditions on the materials in use. The high technology equipment in use today was designed to help us safely explore the more exotic operating environments where the gases no longer behave in an "ideal" manner. This advanced math, including statistics and multivariable calculus, makes possible the solution to such complex dynamic situations as space vehicle reentry. An example is the analysis of the space shuttle reentry pictured to ensure the material properties under this loading condition are appropriate. In this flight regime, the gas is no longer behaving ideally.

Pressure

Main article: Pressure

The symbol used to represent *pressure* in equations is **"p"** or **"P"** with SI units of pascals.

When describing a container of gas, the term pressure (or absolute pressure) refers to the average force per unit area that the gas exerts on the surface of the container. Within this volume, it is sometimes easier to visualize the gas particles moving in straight lines until they collide with the container (see diagram at top of the article). The force imparted by a gas particle into the container during this collision is the change in momentum of the particle.[10] During a collision only the normal (geometry) component of velocity changes. A particle traveling parallel to the wall does not change its momentum. Therefore the average force on a surface must be the average change in linear momentum from all of these gas particle collisions. Pressure is the sum of all the normal components of force exerted by the particles impacting the walls of the container divided by the surface area of the wall.

Temperature

Air balloon shrinks after submersion in liquid nitrogen

Main article: Thermodynamic temperature

The symbol used to represent *temperature* in equations is *T* with SI units of kelvins.

The speed of a gas particle is proportional to its absolute temperature. The volume of the balloon in the video shrinks when the trapped gas particles slow down with the addition of extremely cold nitrogen. The temperature of any physical system is related to the motions of the particles (molecules and atoms) which make up the [gas] system.[11] In statistical mechanics, temperature is the measure of the average kinetic energy stored in a particle. The methods of storing this energy are dictated by the degrees of freedom of the particle itself (energy modes). Kinetic energy added (endothermic process) to gas particles by way of collisions produces linear, rotational, and vibrational motion. In contrast, a molecule in a solid can only increase its vibrational modes with the addition of heat as the lattice crystal structure prevents both linear and rotational motions. These heated gas molecules have a greater speed range which constantly varies due to constant collisions with other particles. The speed range can be described by the Maxwell-Boltzmann distribution. Use of this distribution implies ideal gases near thermodynamic equilibrium for the system of particles being considered.

Specific volume

Main article: Specific volume

The symbol used to represent *specific volume* in equations is **"v"** with SI units of cubic meters per kilogram.

See also: Gas volume

The symbol used to represent **volume** in equations is **"V"** with SI units of cubic meters.

When performing a thermodynamic analysis, it is typical to speak of intensive and extensive properties. Properties which depend on the amount of gas (either by mass or volume) are called *extensive* properties, while properties that do not depend on the amount of gas are called *intensive* properties. **Specific volume** is an example of an *intensive* property because it is the ratio of volume occupied by a *unit of mass* of a gas that is identical throughout a system at equilibrium.[12] 1000 atoms a gas occupy the same space as any other 1000 atoms for any given temperature and pressure. This concept is easier to visualize for solids such as iron which are incompressible compared to gases. Since a gas fills any container in which it is placed, **volume** is an *extensive property*.

Density

Main article: Density

The symbol used to represent **density** in equations is **ρ** (rho) with SI units of kilograms per cubic meter. This term is the reciprocal of specific volume.

Since gas molecules can move freely within a container, their mass is normally characterized by **density**. Density is the amount of mass per unit volume of a substance, or the inverse of specific volume. For gases, the density can vary over a wide range because the particles are free to move closer together when constrained by pressure or volume. This variation of density is referred to as compressibility. Like pressure and temperature, density is a state variable of a gas and the change in density during any process is governed by the laws of thermodynamics. For a static gas, the density is the same throughout the entire container.

Density is therefore a scalar quantity. It can be shown by **kinetic theory** that the density is *inversely* proportional to the size of the container in which a fixed mass of gas is confined. In this case of a fixed mass, the density decreases as the volume increases.

Microscopic

If one could observe a gas under a powerful microscope, one would see a collection of particles (molecules, atoms, ions, electrons, etc.) without any definite shape or volume that are in more or less random motion. These neutral gas particles only change direction when they collide with another particle or with the sides of the container. In an ideal gas, these collisions are perfectly elastic. This particle or microscopic view of a gas is described by the Kinetic-molecular theory. The assumptions behind this theory can be found in the postulates section of Kinetic Theory.

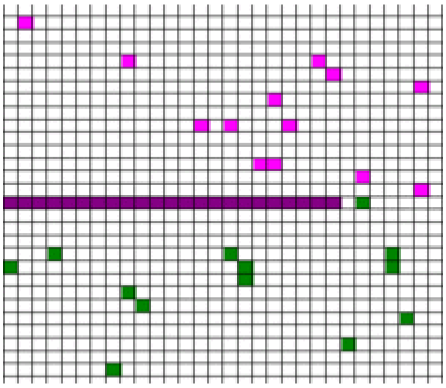
Kinetic theory

Main article: Kinetic theory

Kinetic theory provides insight into the macroscopic properties of gases by considering their molecular composition and motion. Starting with the definitions of momentum and kinetic energy,[13] one can use the conservation of momentum and geometric relationships of a cube to relate macroscopic system properties of temperature and pressure to the microscopic property of kinetic energy per molecule. The theory provides averaged values for these two properties.

The theory also explains how the gas system responds to change. For example, as a gas is heated from absolute zero, when it is (in theory) perfectly still, its internal energy (temperature) is increased. As a gas is heated, the particles speed up and its temperature rises. This results in greater numbers of collisions with the container per unit time due to the higher particle speeds associated with elevated temperatures. The pressure increases in proportion to the number of collisions per unit time.

Brownian motion

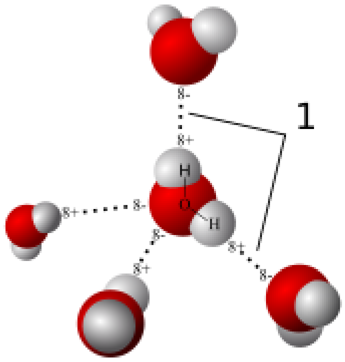


Random motion of gas particles results in diffusion.
Main article: Brownian motion

Brownian motion is the mathematical model used to describe the random movement of particles suspended in a fluid. The gas particle animation, using pink and green particles, illustrates how this behavior results in the spreading out of gases (entropy). These events are also described by particle theory.

Since it is at the limit of (or beyond) current technology to observe individual gas particles (atoms or molecules), only theoretical calculations give suggestions about how they move, but their motion is different from Brownian motion because Brownian motion involves a smooth drag due to the frictional force of many gas molecules, punctuated by violent collisions of an individual (or several) gas molecule(s) with the particle. The particle (generally consisting of millions or billions of atoms) thus moves in a jagged course, yet not so jagged as would be expected if an individual gas molecule were examined.

Intermolecular forces



When gases are compressed, intermolecular forces like those shown here start to play a more active role.
Main articles: van der Waals force and Intermolecular force

As discussed earlier, momentary attractions (or repulsions) between particles have an effect on gas dynamics. In physical chemistry, the name given to these intermolecular forces is *van der Waals force*. These forces play a key role in determining physical properties of a gas such as viscosity and flow rate (see physical characteristics section). Ignoring these forces in certain conditions (see Kinetic-molecular theory) allows a real gas to be treated like an ideal gas. This assumption allows the use of ideal gas laws which greatly simplifies calculations.

Proper use of these gas relationships requires the Kinetic-molecular theory (KMT). When gas particles possess a magnetic charge or Intermolecular force they gradually influence one another as the spacing between them is reduced (the hydrogen bond model illustrates one example). In the absence of any charge, at some point when the spacing between gas particles is greatly reduced they can no longer avoid collisions between themselves at normal gas temperatures. Another case for increased collisions among gas particles would include a fixed volume of gas, which upon heating would contain very fast particles. *This means that these ideal equations provide reasonable results* except *for extremely high pressure (compressible) or high temperature (ionized) conditions.* Notice that all of these excepted conditions allow energy transfer to take place within the gas system. The absence of these internal transfers is what is referred to as ideal conditions in which the energy exchange occurs only at the boundaries of the system. Real gases experience some of these collisions and intermolecular forces. When these collisions are statistically negligible (incompressible), results from these ideal equations are still meaningful. If the gas particles are compressed into close proximity they behave more like a liquid (see fluid dynamics).

Simplified models

Main article: Equation of state

An *equation of state* (for gases) is a mathematical model used to roughly describe or predict the state properties of a gas. At present, there is no single equation of state that accurately predicts the properties of all gases under all conditions. Therefore, a number of much more accurate equations of state have been developed for gases in specific temperature and pressure ranges. The "gas models" that are most widely discussed are "perfect gas", "ideal gas" and "real gas". Each of these models has its own set of assumptions to facilitate the analysis of a given thermodynamic system.[14] Each successive model expands the temperature range of coverage to which it applies.

Ideal and perfect gas models

Main article: Perfect gas

The equation of state for an ideal or perfect gas is the ideal gas law and reads

$$PV = nRT,$$

where P is the pressure, V is the volume, n is amount of gas (in mol units), R is the universal gas constant, 8.314 J/(mol K), and T is the temperature. Written this way, it is sometimes called the "chemist's version", since it emphasizes the number of molecules n. It can also be written as

$$P = \rho R_s T,$$

where R_s is the specific gas constant for a particular gas, in units J/(kg K), and $\rho = m/V$ is density. This notation is the "gas dynamicist's" version, which is more practical in modeling of gas flows involving acceleration without chemical reactions.

The ideal gas law does not make an assumption about the specific heat of a gas. In the most general case, the specific heat is a function of both temperature and pressure. If the pressure-dependence is neglected (and possibly the temperature-dependence as well) in a particular application, sometimes the gas is said to be a perfect gas, although the exact assumptions may vary depending on the author and/or field of science.

For an ideal gas, the ideal gas law applies without restrictions on the specific heat. An ideal gas is a simplified "real gas" with the assumption that the compressibility factor Z is set to 1 meaning that this pneumatic ratio remains constant. A compressibility factor of one also requires the four state variables to follow the ideal gas law.

This approximation is more suitable for applications in engineering although simpler models can be used to produce a "ball-park" range as to where the real solution should lie. An example where the "ideal gas approximation" would be suitable would be inside a combustion chamber of a jet engine.[15] It may also be useful to keep the elementary reactions and chemical dissociations for calculating emissions.

Real gas

21 April 1990 eruption of Mount Redoubt, Alaska, illustrating real gases not in thermodynamic equilibrium.

Main article: Real gas

Each one of the assumptions listed below adds to the complexity of the problem's solution. As the density of a gas increases with pressure rises, the intermolecular forces play a more substantial role in gas behavior which results in the ideal gas law no longer providing "reasonable" results. At the upper end of the engine temperature ranges (e.g. combustor sections – 1300 K), the complex fuel particles absorb internal energy by means of rotations and vibrations that cause their specific heats to vary from those of diatomic molecules and noble gases. At more than double that temperature, electronic excitation and dissociation of the gas particles begins to occur causing the pressure to adjust to a greater number of particles (transition from gas to plasma).[16] Finally, all of the thermodynamic processes were presumed to describe uniform gases whose velocities varied according to a fixed distribution. Using a non-equilibrium situation implies the flow field must be characterized in some manner to enable a solution. One of the first attempts to expand the boundaries of the ideal gas law was to include coverage for different thermodynamic processes by adjusting the equation to read $pV^n = constant$ and then varying the n through different values such as the specific heat ratio, γ.

Real gas effects include those adjustments made to account for a greater range of gas behavior:

- Compressibility effects (Z allowed to vary from 1.0)
- Variable heat capacity (specific heats vary with temperature)
- Van der Waals forces (related to compressibility, can substitute other equations of state)
- Non-equilibrium thermodynamic effects
- Issues with molecular dissociation and elementary reactions with variable composition.

For most applications, such a detailed analysis is excessive. Examples where "Real Gas effects" would have a significant impact would be on the Space Shuttle re-entry where extremely high temperatures and pressures are present or the gases produced during geological events as in the image of the 1990 eruption of Mount Redoubt.

Historical synthesis

Boyle's law

See also: Gas laws

Boyle's equipment.

Main article: Boyle's law

> Boyle's Law was perhaps the first expression of an equation of state. In 1662 Robert Boyle performed a series of experiments employing a J-shaped glass tube, which was sealed on one end. Mercury was added to the tube, trapping a fixed quantity of air in the short, sealed end of the tube. Then the volume of gas was carefully measured as additional mercury was added to the tube. The pressure of the gas could be determined by the difference between the mercury level in the short end of the tube and that in the long, open end. Through these experiments, Boyle noted that the *gas volume varied inversely with the pressure*.[17] The image of Boyle's Equipment shows some of the exotic tools used by Boyle during his study of gases.

- **Boyle's Law** – describes a gas in which the number of particles and Temperature are constant.
- ***PV = constant*** in this situation *constant* = *nRT* from the ideal gas law.

Law of volumes

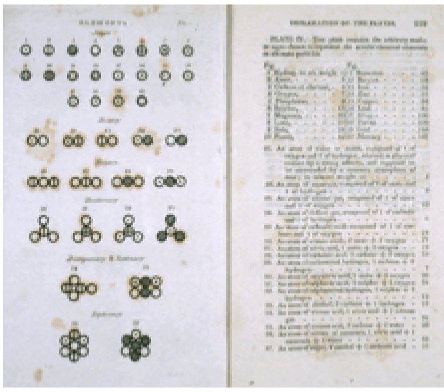
Main articles: Charles's law and Gay-Lussac's Law

In 1787, the French physicist and balloon pioneer, Jacques Charles, found that oxygen, nitrogen, hydrogen, carbon dioxide, and air expand to the same extent over the same 80 kelvin interval.

In 1802, Joseph Louis Gay-Lussac published results of similar, though more extensive experiments, indicating a linear relationship between volume and temperature. Gay-Lussac credited Charle's earlier work by naming the law in his honor. In the absence of this linkage, Dalton could have been in contention for this honor for his previously published work on partial pressures.

- **Law of Volumes** – Both Charles and Gay-Lussac played a role in developing this relationship.[18]
- **$V/T = constant$** – notice that $constant = nR/P$ from the ideal gas law.

Avogadro's law



Dalton's notation.

Main article: Avogadro's law

In 1811, Amedeo Avogadro verified that equal volumes of pure gases contain the same number of particles. His theory was not generally accepted until 1858 when another Italian chemist Stanislao Cannizzaro was able to explain non-ideal exceptions. For his work with gases a century prior, the number that bears his name Avogadro's constant represents the number of atoms found in 12 grams of elemental carbon-12 ($6.022\times10^{23}\ \text{mol}^{-1}$). This specific number of gas particles, at standard temperature and pressure (ideal gas law) occupies 22.40 liters, which is referred to as the molar volume.

- **Avogadro's Law** – describes a gas in a container in which the pressure and temperature are constant. The simplified form for the ideal gas law follows:
- **$V/n = constant$** – notice that $constant = RT/P$ from the ideal gas law.

Dalton's law

Main article: Dalton's law

In 1801, John Dalton published the **Law of Partial Pressures** from his work with ideal gas law relationship: The pressure of a mixture of gases is equal to the sum of the pressures of all of the constituent gases alone. Mathematically, this can be represented for n species as:

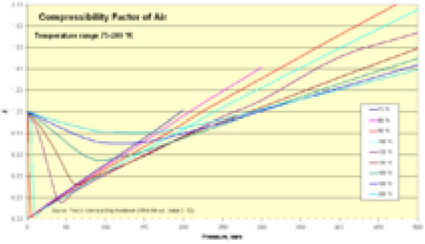
$$Pressure_{total} = Pressure_1 + Pressure_2 + \ldots + Pressure_n$$

The image of Dalton's journal depicts symbology he used as shorthand to record the path he followed. Among his key journal observations upon mixing unreactive "elastic fluids" (gases) were the following.[19]:

- Unlike liquids, heavier gases did not drift to the bottom upon mixing.
- Gas particle identity played no role in determining final pressure (they behaved as if their size was negligible).

Special topics

Compressibility



Compressibility factors for air.
Main article: Compressibility factor

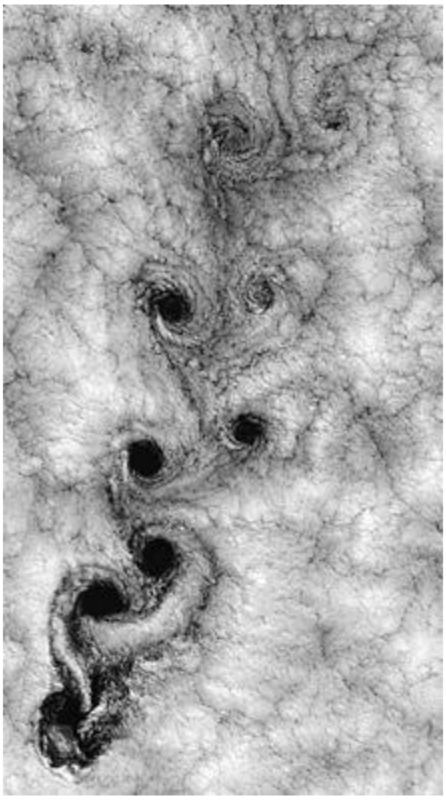
Thermodynamicists use this factor (*Z*) to alter the ideal gas equation to account for compressibility effects of real gases. This factor represents the ratio of actual to ideal specific volumes. It is sometimes referred to as a "fudge-factor" or correction to expand the useful range of the ideal gas law for design purposes. *Usually* this *Z* value is very close to unity. The compressibility factor image illustrates how Z varies over a range of very cold temperatures.

Reynolds number

Main article: Reynolds number

In fluid mechanics, the Reynolds number is the ratio of inertial forces ($v_s\rho$) to viscous forces (μ/L). It is one of the most important dimensionless numbers in fluid dynamics and is used, usually along with other dimensionless numbers, to provide a criterion for determining dynamic similitude. As such, the Reynolds number provides the link between modeling results (design) and the full-scale actual conditions. It can also be used to characterize the flow.

Viscosity

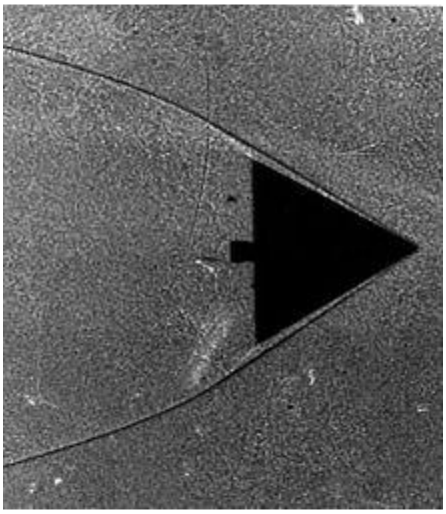


Satellite view of weather pattern in vicinity of Robinson Crusoe Islands on 15 September 1999, shows a unique turbulent cloud pattern called a Kármán vortex street

Main article: Viscosity

Viscosity, a physical property, is a measure of how well adjacent molecules stick to one another. A solid can withstand a shearing force due to the strength of these sticky intermolecular forces. A fluid will continuously deform when subjected to a similar load. While a gas has a lower value of viscosity than a liquid, it is still an observable property. If gases had no viscosity, then they would not stick to the surface of a wing and form a boundary layer. A study of the delta wing in the Schlieren image reveals that the gas particles stick to one another (see Boundary layer section).

Turbulence



Delta wing in wind tunnel. The shadows form as the indices of refraction change within the gas as it compresses on the leading edge of this wing.

Main article: Turbulence

In fluid dynamics, **turbulence** or turbulent flow is a flow regime characterized by chaotic, stochastic property changes. This includes low momentum diffusion, high momentum convection, and rapid variation of pressure and velocity in space and time. The Satellite view of weather around Robinson Crusoe Islands illustrates just one example.

Boundary layer

Main article: Boundary layer

Particles will, in effect, "stick" to the surface of an object moving through it. This layer of particles is called the **boundary layer**. At the surface of the object, it is essentially static due to the friction of the surface. The object, with its boundary layer is effectively the new shape of the object that the rest of the molecules "see" as the object approaches. This boundary layer *can* separate from the surface, essentially creating a new surface and completely changing the flow path. The classical example of this is a stalling airfoil. The delta wing image clearly shows the boundary layer thickening as the gas flows from right to left along the leading edge.

Maximum entropy principle

Main article: Principle of maximum entropy

As the total number of degrees of freedom approaches infinity, the system will be found in the macrostate that corresponds to the highest multiplicity. In order to illustrate this principle, observe the skin temperature of a frozen metal bar. Using a thermal image of the skin temperature, note the temperature distribution on the surface. This initial observation of temperature represents a "microstate." At some future time, a second observation of the skin temperature produces a second microstate. By continuing this observation process, it is possible to produce a series of microstates that illustrate the thermal history of the bar's surface. Characterization of this historical series of microstates is possible by choosing the macrostate that successfully classifies them all into a single grouping.

Thermodynamic equilibrium

Main article: Thermodynamic equilibrium

When energy transfer ceases from a system, this condition is referred to as thermodynamic equilibrium. Usually this condition implies the system and surroundings are at the same temperature so that heat no longer transfers between them. It also implies that external forces are balanced (volume does not change), and all chemical reactions within the system are complete. The timeline varies for these events depending on the system in question. A container of ice allowed to melt at room temperature takes hours, while in semiconductors the heat transfer that occurs in the device transition from an on to off state could be on the order of a few nanoseconds.

See also

- Air conditioning
- Argon
- Atmosphere
- Bellows
- Breath
- Carbon Dioxide
- Cavendish, Henry
- Chlorine
- Clouds
- Cooking
- Electrolysis
- Flame tests
- Kites
- Lavoisier, Antoine
- Lift (soaring)
- Lighting
- Lightning
- List of plasma (physics) articles
- Liquid
- Lung
- Mixmaster dynamics
- Nitrogen
- Odor
- Oxygen
- Parachute
- Petroleum
- Plasma (physics)
- Priestley, Joseph
- Ramsay, William
- Sailing
- Solid
- Thoracic diaphragm
- Troposphere
- Turbines
- Vapor
- Volcanic gas
- Weather
- Wind
- Wind mill
- Wind turbine

Notes

1. ^ This early 20th century discussion infers what is regarded as the plasma state. See page 137 of American Chemical Society, Faraday Society, Chemical Society (Great Britain) *The Journal of physical chemistry, Volume 11* Cornell (1907).

2. ^ The work by T. Zelevinski provides another link to latest research about Strontium in this new field of study. See Tanya Zelevinsky (2009). "84Sr—just right for forming a Bose-Einstein condensate". *Physics* **2**: 94. http://physics.aps.org/articles/v2/94.
3. ^ for links material on the Bose-Einstein condensate see Quantum Gas Microscope Offers Glimpse Of Quirky Ultracold Atoms. ScienceDaily. 4 November 2009.
4. ^ J. B. van Helmont, *Ortus medicinae*. … (Amsterdam, (Netherlands): Louis Elzevir, 1652 (first edition: 1648)). The word "gas" first appears on page 58, where he mentions: "… Gas (meum scil. inventum) …" (… gas (namely, my discovery) …). On page 59, he states: "… in nominis egestate, halitum illum, Gas vocavi, non longe a Chao …" (… in need of a name, I called this vapor "gas", not far from "chaos" …)
5. ^ Harper, Douglas. "gas". *Online Etymology Dictionary*. http://www.etymonline.com/index.php?term=gas.
6. ^ The authors make the connection between molecular forces of metals and their corresponding physical properties. By extension, this concept would apply to gases as well, though not universally. Cornell (1907) pp. 164–5.
7. ^ One noticeable exception to this physical property connection is conductivity which varies depending on the state of matter (ionic compounds in water) as described by Michael Faraday in the 1833 when he noted that ice does not conduct a current. See page 45 of John Tyndall's *Faraday as a Discoverer* (1868).
8. ^ John S. Hutchinson (2008). *Concept Development Studies in Chemistry*. p. 67. http://cnx.org/content/col10264/latest/.
9. ^ Anderson, p.501
10. ^ J. Clerk Maxwell (1904). *Theory of Heat*. Mineola: Dover Publications. pp. 319–20. ISBN 0-486-41735-2.
11. ^ See pages 137–8 of Society, Cornell (1907).
12. ^ Kenneth Wark (1977). *Thermodynamics* (3 ed.). McGraw-Hill. p. 12. ISBN 0-07-068280-1.
13. ^ For assumptions of Kinetic Theory see McPherson, pp.60–61
14. ^ Anderson, pp. 289–291
15. ^ John, p.205
16. ^ John, pp. 247–56
17. ^ McPherson, pp.52–55
18. ^ McPherson, pp.55–60
19. ^ John P. Millington (1906). *John Dalton*. pp. 72, 77–78.

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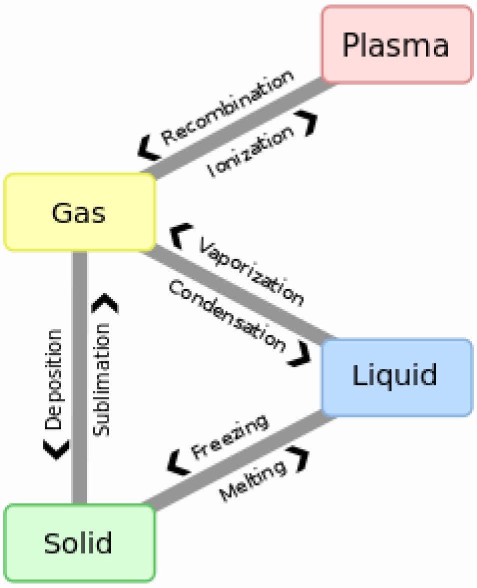
States of matter (classic)

State

- Solid ·
- Liquid ·
- **Gas** / Vapor ·
- Plasma

Low energy

- Bose–Einstein condensate ·
- Fermionic condensate ·
- Degenerate matter ·
- Quantum Hall ·
- Rydberg matter ·
- Strange matter ·
- Superfluid ·
- Supersolid



High energy

- QCD matter ·
- Quark–gluon plasma ·
- Supercritical fluid

Other states

- Colloid ·
- Glass ·
- Liquid crystal ·
- Magnetically ordered

  (

  - Antiferromagnet ·
  - Ferrimagnet ·

	○ Ferromagnet) · • String-net liquid · • Superglass
Transitions	• Boiling · • Boiling point · • Condensation · • Critical line · • Critical point · • Crystallization · • Deposition · • Evaporation · • Flash evaporation · • Freezing · • Ionization · • Lambda point · • Melting · • Melting point · • Recombination · • Regelation · • Saturated fluid · • Sublimation · • Supercooling · • Triple point · • Vaporization · • Vitrification

Quantities	Enthalpy of fusion · Enthalpy of sublimation · Enthalpy of vaporization · Latent heat · Latent internal energy · Trouton's ratio · Volatility

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