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hydrogen

chemical element

Alternate titles: *H*Written by [William Lee Jolly](#)Fact-checked by [The Editors of Encyclopaedia Britannica](#)Last Updated: [Article History](#)

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Summary

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hydrogen (H), a colourless, odourless, tasteless, flammable gaseous substance that is the simplest member of the family of chemical elements. The hydrogen atom has a nucleus consisting of a [proton bearing](#) one unit of positive electrical charge; an electron, bearing one unit of negative [electrical charge](#), is also associated with this nucleus. Under ordinary conditions, hydrogen gas is a loose aggregation of hydrogen molecules, each consisting of a pair of atoms, a [diatomic molecule](#), H₂. The earliest known important chemical property of hydrogen is that it burns with [oxygen](#) to form [water](#), H₂O; indeed, the name hydrogen is derived from Greek words meaning “maker of water.”

Although hydrogen is the most abundant element in the [universe](#) (three times as abundant as [helium](#), the next most widely occurring element), it makes up only about 0.14 percent of Earth's crust by weight. It occurs, however, in vast quantities as part of the water in oceans, [ice](#) packs, rivers, lakes, and the atmosphere. As part of innumerable [carbon compounds](#), hydrogen is present in all animal and vegetable tissue and in petroleum. Even though it is often said that there are more known compounds of carbon than of any other element, the fact is that, since hydrogen is contained in almost all carbon compounds

gen

atomic number: 1 [1.00784, 1.00811] atomic weight

symbol: H acid-base properties of higher-valence oxides

electron configuration: 1s¹ crystal structure

name: hydrogen physical state at 20 °C (68 °F)

Other nonmetals Gas

Hexagonal Equal relative strength

chemical properties of hydrogen

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Key People: [Henry Cavendish](#) • [Antoine Lavoisier](#) • [Otto Struve](#) • [Anders Jonas Ångström](#) • [Fritz Wolfgang London](#)

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and also forms a multitude of compounds with all other elements (except some of the noble gases), it is possible that hydrogen compounds are more numerous.

Elementary hydrogen finds its principal industrial application in the manufacture of [ammonia](#) (a [compound](#) of hydrogen and [nitrogen](#), NH_3) and in the [hydrogenation](#) of [carbon monoxide](#) and organic compounds.

Hydrogen has three known isotopes. The mass numbers of hydrogen's isotopes are 1, 2, and 3, the most abundant being the mass 1 [isotope](#) generally called hydrogen (symbol H, or ^1H) but also known as [protium](#). The mass 2 isotope, which has a nucleus of one proton and one [neutron](#) and has been named [deuterium](#), or heavy hydrogen (symbol D, or ^2H), [constitutes](#) 0.0156 percent of the ordinary mixture of hydrogen. Tritium (symbol T, or ^3H), with one proton and two neutrons in each nucleus, is the mass 3 isotope and constitutes about 10^{-15} to 10^{-16} percent of hydrogen. The practice of giving distinct names to the hydrogen isotopes is justified by the fact that there are significant differences in their properties.



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Paracelsus, physician and alchemist, in the 16th century unknowingly experimented with hydrogen when he found that a flammable gas was evolved when a [metal](#) was dissolved in [acid](#). The gas, however, was confused with other flammable gases, such as hydrocarbons and carbon monoxide. In 1766 [Henry Cavendish](#), English chemist and physicist, showed that hydrogen, then called flammable [air](#), [phlogiston](#), or the flammable principle, was distinct from other combustible gases because of its [density](#) and the amount of it that evolved from a given amount of acid and metal. In 1781 Cavendish confirmed previous observations that water was formed when hydrogen was burned, and [Antoine-Laurent Lavoisier](#), the father of modern [chemistry](#), coined the French word *hydrogène* from which the English form is [derived](#). In 1929 Karl Friedrich Bonhoeffer, a German physical chemist, and Paul Harteck, an Austrian chemist, on the basis of earlier theoretical work, showed that ordinary hydrogen is a mixture of two kinds of molecules, *ortho*-hydrogen and *para*-hydrogen. Because of the simple structure of hydrogen, its

properties can be theoretically calculated relatively easily. Hence hydrogen is often used as a theoretical model for more complex atoms, and the results are applied qualitatively to other atoms.

Physical and chemical properties

The Table lists the important properties of molecular hydrogen, H₂. The extremely low melting and boiling points result from weak forces of attraction between the molecules. The existence of these weak intermolecular forces is also revealed by the fact that, when hydrogen gas expands from high to [low pressure](#) at room temperature, its temperature rises, whereas the temperature of most other gases falls. According to thermodynamic principles, this implies that repulsive forces [exceed](#) attractive forces between hydrogen molecules at room temperature—otherwise, the expansion would cool the hydrogen. In fact, at -68.6° C attractive forces predominate, and hydrogen, therefore, cools upon being allowed to expand below that temperature. The cooling effect becomes so pronounced at temperatures below that of liquid nitrogen (-196° C) that the effect is utilized to achieve the liquefaction temperature of hydrogen gas itself.

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Some properties of normal hydrogen and deuterium		
	normal hydrogen	deuterium
Atomic hydrogen		
atomic number	1	1
atomic weight	1.0080	2.0141
ionization potential	13.595 electron volts	13.600 electron volts
electron affinity	0.7542 electron volts	0.754 electron volts
nuclear spin	1/2	1
nuclear magnetic moment (nuclear magnetons)	2.7927	0.8574
nuclear quadrupole moment	0	2.77(10 ⁻²⁷) square centimetres
electronegativity (Pauling)	2.1	~2.1
Molecular hydrogen		
bond distance	0.7416 angstrom	0.7416 angstrom
dissociation energy (25 degrees C)	104.19 kilocalories per mole	105.97 kilocalories per mole
ionization potential	15.427 electron volts	15.457 electron volts
density of solid	0.08671 gram per cubic centimetre	0.1967 gram per cubic centimetre
melting point	-259.20 degrees Celsius	-254.43 degrees Celsius
heat of fusion	28 calories per mole	47 calories per mole
density of liquid	0.07099 (-252.78 degrees)	0.1630 (-249.75 degrees)
boiling point	-252.77 degrees Celsius	-249.49 degrees Celsius
heat of vaporization	216 calories per mole	293 calories per mole
critical temperature	-240.0 degrees Celsius	-243.8 degrees Celsius
critical pressure	13.0 atmospheres	16.4 atmospheres
critical density	0.0310 gram per cubic centimetre	0.0668 gram per cubic centimetre
heat of combustion to water (g)	-57.796 kilocalories per mole	-59.564 kilocalories per mole

Hydrogen is transparent to visible light, to infrared light, and to [ultraviolet light](#) to wavelengths below 1800 Å. Because its [molecular weight](#) is lower than that of any other gas, its molecules have a velocity higher than those of any other gas at a given temperature and it [diffuses](#) faster than any other gas. Consequently, [kinetic energy](#) is distributed faster through hydrogen than through any other gas; it has, for example, the greatest heat conductivity.

A **molecule** of hydrogen is the simplest possible molecule. It consists of two protons and two electrons held together by electrostatic forces. Like atomic hydrogen, the assemblage can exist in a number of energy levels.

Ortho-hydrogen and para-hydrogen

Two types of molecular hydrogen (*ortho* and *para*) are known. These differ in the magnetic **interactions** of the **protons** due to the **spinning** motions of the protons. In *ortho*-hydrogen, the spins of both protons are aligned in the same direction—that is, they are parallel. In *para*-hydrogen, the spins are aligned in opposite directions and are therefore antiparallel. The relationship of **spin** alignments determines the magnetic properties of the **atoms**. Normally, transformations of one type into the other (*i.e.*, conversions between *ortho* and *para* molecules) do not occur and *ortho*-hydrogen and *para*-hydrogen can be regarded as two distinct modifications of hydrogen. The two forms may, however, interconvert under certain conditions. **Equilibrium** between the two forms can be established in several ways. One of these is by the introduction of **catalysts** (such as activated charcoal or various paramagnetic substances); another method is to apply an electrical discharge to the gas or to heat it to a high temperature.

The concentration of *para*-hydrogen in a mixture that has achieved **equilibrium** between the two forms depends on the temperature as shown by the following figures:

– 253.1° C	99.82%	– 153.1° C	32.87%
– 223.1° C	76.89%	0° C	25.13 %
– 193.1° C	48.39%	200° C	25.00%

Essentially pure *para*-hydrogen can be produced by bringing the mixture into contact with charcoal at the temperature of liquid hydrogen; this converts all the *ortho*-hydrogen into *para*-hydrogen. The *ortho*-hydrogen, on the other hand, cannot be prepared directly from the mixture because the concentration of *para*-hydrogen is never less than 25 percent.

The two forms of hydrogen have slightly different physical properties. The **melting point** of *para*-hydrogen is 0.10° lower than that of a 3:1 mixture of *ortho*-hydrogen and *para*-hydrogen. At –252.77° C the pressure exerted by the vapour over liquid *para*-hydrogen is 1.035 atmospheres (one atmosphere is the pressure of the atmosphere at **sea level** under standard conditions, equal to about 14.69 pounds per square inch), compared with 1.000 atmosphere for the **vapour pressure** of the 3:1 *ortho*–*para* mixture. As a result of the different vapour pressures of *para*-hydrogen and *ortho*-hydrogen, these forms of

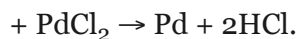
hydrogen can be separated by low-temperature [gas chromatography](#), an [analytical](#) process that separates different atomic and molecular species on the basis of their differing volatilities.

Reactivity of hydrogen

One [molecule](#) of hydrogen dissociates into two atoms ($\text{H}_2 \rightarrow 2\text{H}$) when an energy equal to or greater than the dissociation energy (*i.e.*, the amount of energy required to break the [bond](#) that holds together the atoms in the molecule) is supplied. The dissociation energy of molecular hydrogen is 104,000 calories per mole—written 104 kcal/mole (mole: the [molecular weight](#) expressed in grams, which is two grams in the case of hydrogen). [Sufficient](#) energy is obtained, for example, when the [gas](#) is brought into contact with a white-hot [tungsten](#) filament or when an electric discharge is established in the gas. If atomic hydrogen is generated in a system at [low pressure](#), the atoms will have a significant lifetime—*e.g.*, 0.3 second at a pressure of 0.5 millimetre of [mercury](#). Atomic hydrogen is very reactive. It combines with most elements to form hydrides (*e.g.*, [sodium hydride](#), NaH), and it reduces metallic oxides, a reaction that produces the [metal](#) in its elemental state. The surfaces of metals that do not combine with hydrogen to form stable hydrides (*e.g.*, platinum) catalyze the recombination of hydrogen atoms to form hydrogen molecules and are thereby heated to incandescence by the energy that this reaction releases.

Molecular hydrogen can react with many elements and [compounds](#), but at room temperature the [reaction rates](#) are usually so low as to be negligible. This apparent inertness is in part related to the very high dissociation energy of the molecule. At elevated temperatures, however, the reaction rates are high.

Sparks or certain radiations can cause a mixture of hydrogen and [chlorine](#) to react explosively to yield [hydrogen chloride](#), as represented by the equation $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$. Mixtures of hydrogen and [oxygen](#) react at a measurable rate only above 300° C, according to the equation $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$. Such mixtures containing 4 to 94 percent hydrogen ignite when heated to 550°–600° C or when brought into contact with a [catalyst](#), spark, or flame. The explosion of a 2:1 mixture of hydrogen and oxygen is especially violent. Almost all metals and nonmetals react with hydrogen at high temperatures. At elevated temperatures and pressures hydrogen [reduces](#) the oxides of most metals and many metallic salts to the metals. For example, hydrogen gas and ferrous oxide react, yielding metallic [iron](#) and [water](#), $\text{H}_2 + \text{FeO} \rightarrow \text{Fe} + \text{H}_2\text{O}$; hydrogen gas reduces [palladium](#) chloride to form palladium metal and hydrogen chloride, H_2



Hydrogen is absorbed at high temperatures by many [transition metals](#) (scandium, 21, through [copper](#), 29; [yttrium](#), 39, through [silver](#), 47; [hafnium](#), 72, through [gold](#), 79); and metals of the actinoid (actinium, 89, through [lawrencium](#), 103) and [lanthanoid](#) series (lanthanum, 57, through [lutetium](#), 71) to form hard, [alloy-like](#) hydrides. These are often called [interstitial](#) hydrides because, in many cases, the metallic [crystal lattice](#) merely expands to [accommodate](#) the dissolved hydrogen without any other change.

Hydrogen bond

Some covalently bonded hydrides have a hydrogen [atom](#) bound simultaneously to two separate electronegative atoms, which are then said to be hydrogen bonded. The strongest hydrogen bonds involve the small, highly electronegative atoms of [fluorine](#) (F), oxygen, and [nitrogen](#). In the bifluoride ion, HF_2^- , the hydrogen atom links two fluorine atoms. In the crystal structure of ice, each oxygen atom is

The helical (spiral) configurations of certain enormous molecular chains, as in proteins, are held together by hydrogen bonds. Extensive hydrogen bonding in the [liquid state](#) explains why hydrogen fluoride (HF), water (H_2O), and [ammonia](#) (NH_3) have boiling points much higher than those of their heavier [analogues](#), hydrogen chloride (HCl), [hydrogen sulfide](#) (H_2S), and [phosphine](#) (PH_3). Thermal energy required to break up the hydrogen bonds and to permit vaporization is available only at the higher boiling temperatures.

The hydrogen in a strong [acid](#), such as hydrochloric (HCl) or nitric (HNO_3), behaves quite differently. When these acids dissolve in water, [hydrogen](#) in the form of a [proton](#), H^+ , separates completely from the negatively charged ion, the anion (Cl^- or NO_3^-), and [interacts](#) with the water molecules. The proton is strongly attached to one water molecule (hydrated) to form the oxonium [ion](#) (H_3O^+ , sometimes called hydronium ion), which in turn is hydrogen-bonded to other water molecules, forming species with formulas such as $\text{H}(\text{H}_2\text{O})_n^+$ (the subscript n indicates the number of H_2O molecules involved). The

reduction of H^+ (reduction is the chemical change in which an atom or ion gains one or more electrons) can be represented as the half reaction: $\text{H}^+ + e^- \rightarrow \frac{1}{2}\text{H}_2$. The energy needed to bring about this reaction can be expressed as a reduction potential. The reduction potential for hydrogen is taken by convention to be zero, and all metals with negative reduction potentials—*i.e.*, metals that are less easily reduced (more easily oxidized; *e.g.*, zinc: $\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}$, -0.763 volt)—can, in principle, displace hydrogen from a strong acid solution: $\text{Zn} + 2\text{H}^+ \rightarrow \text{Zn}^{2+} + \text{H}_2$. Metals with positive reduction potentials (*e.g.*, silver: $\text{Ag}^+ + e^- \rightarrow \text{Ag}$, $+0.7995$ volt) are inert toward the aqueous hydrogen ion.

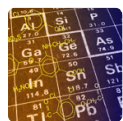
Isotopes of hydrogen

By means of the mass spectrograph he had invented, [Francis William Aston](#) in 1927 observed that the line for hydrogen corresponded to an [atomic weight](#) on the chemical scale of 1.00756. This value differed by more than the probable experimental error from the value based on the combining weights of hydrogen [compounds](#), 1.00777. Other workers showed that the discrepancy could be removed by postulating the existence of a hydrogen [isotope](#) of mass 2 in the proportion of one [atom](#) of ^2H (or D) to 4,500 atoms of ^1H . The problem interested the U.S. chemist [Harold C. Urey](#), who from theoretical principles predicted a difference in the vapour pressures of hydrogen (H_2) and hydrogen deuteride (HD) and thus the possibility of [separating](#) these substances by [distillation](#) of liquid hydrogen. In 1931 Urey and two collaborators detected [deuterium](#) by its atomic spectrum in the residue of a distillation of liquid hydrogen. Deuterium was first prepared in pure form by the electrolytic method of concentration: when a water solution of an electrolyte, such as [sodium hydroxide](#), is [electrolyzed](#), the hydrogen formed at the [cathode](#) contains a smaller fraction of deuterium than the water, and thus deuterium is concentrated in the residue. Almost pure deuterium oxide (D_2O , [heavy water](#)) is obtained when the solution is reduced to 0.00001 of its original volume. Deuterium can be concentrated also by the fractional distillation of [water](#) and by various chemical exchange reactions such as the following (g and l indicate gaseous and liquid states, respectively): $\text{H}_2\text{O}(\text{g}) + \text{HD}(\text{g}) \rightleftharpoons \text{HDO}(\text{g}) + \text{H}_2(\text{g})$; $\text{HDO}(\text{g}) + \text{H}_2\text{S}(\text{g}) \rightleftharpoons \text{HDS}(\text{g}) + \text{H}_2\text{O}(\text{g})$; $\text{NH}_3(\text{l}) + \text{HD}(\text{g}) \rightleftharpoons \text{NH}_2\text{D}(\text{l}) + \text{H}_2(\text{g})$.

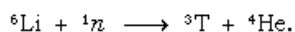
[Tritium](#) (T) was first prepared in 1935 by bombarding deuterium (in the form of deuterophosphoric acid) with high-energy deuterons (deuterium nuclei):



Tritium is present in minute concentrations in natural water. It is formed continuously in the upper atmosphere by cosmic-ray-induced nuclear reactions. [Cosmic rays](#), consisting mainly of high-energy protons, react with [nitrogen](#) atoms to form neutrons, which in turn react with more nitrogen atoms to form tritium:


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118 Names and Symbols of the Periodic Table Quiz


This naturally formed tritium ends up in the form of water and reaches the surface of Earth in rain. Tritium is **radioactive**; it has a half-life of 12.5 years, **decaying** to a very soft (low energy) negative **beta particle** (electron; the positive beta particle is called a positron) and a helium-3 nucleus. When a sample of water is stored, it gradually loses its tritium because of **radioactive** decay. Thus by analyzing water for its tritium content, it is possible to elucidate details of water circulation among oceans, the atmosphere, rivers, and lakes. Tritium is made artificially in nuclear reactors by the reaction of thermal neutrons with lithium:



Corresponding compounds of the hydrogen isotopes differ slightly in their physical properties. This difference is shown by the properties of the waters, listed in the Table, and of the elements, listed in the following Table. The same is true of their chemical properties, both thermodynamic and **kinetic**. Both deuterium and tritium are useful as **isotopic tracers** for the investigation of chemical structures and of reaction mechanisms. Generally the value of a tracer arises from the fact that, although its difference in **mass** or its radioactivity permits its detection, it is essentially active in the same way that the ordinary atoms of the element are. For most elements, a change of one or of a few mass units is such a small percentage of the total mass that the chemical differences between **isotopes** are negligible. For hydrogen, however, chemical reactions involving the different isotopes proceed at measurably different rates. These kinetic-isotope effects can be **utilized** in detailed studies of reaction mechanisms. The rates of reactions of compounds containing deuterium or tritium are usually less than those of the corresponding compounds of ordinary hydrogen.

Physical properties of the waters			
	hydrogen oxide	deuterium oxide	tritium oxide
density at 25 degrees Celsius in grams per millilitre	0.99707	1.10451	—
melting point, degrees Celsius	0	3.81	4.49
boiling point, degrees Celsius	100	101.41	—
temperature of maximum density, degrees Celsius	3.98	11.21	13.4
maximum density in grams per millilitre	1.00000	1.10589	1.21502

The replacement of hydrogen by deuterium in biological systems can markedly alter the delicately

balanced processes. It has been established that neither plants nor animals continue to live and thrive in water containing deuterium oxide in high concentrations.

Deuterium and tritium are of interest in connection with thermonuclear (fusion) reactions. The explosion of a [hydrogen bomb](#) involves the collision and [fusion](#) of light nuclei, including deuterium and tritium. Should a method be found for controlling such fusion processes, as was [done](#) with the fission process of the earlier [atomic bomb](#), the raw material for a practically unlimited supply of energy would be available in the deuterium content of water. Such fusion reactions are the source of [solar energy](#).

Deuterium oxide is useful in nuclear reactors as a moderator to slow down but not appreciably capture neutrons. It has the advantage of being a liquid that absorbs neutrons only slightly.

Production and applications of hydrogen

The most important [industrial](#) method for the production of hydrogen is the catalytic steam–hydrocarbon process, in which gaseous or vaporized [hydrocarbons](#) are treated with [steam](#) at high pressure over a [nickel catalyst](#) at 650°–950° C to produce [carbon](#) oxides and hydrogen: $C_nH_{2n+2} + nH_2O \rightarrow nCO + (2n + 1)H_2$; $C_nH_{2n+2} + 2nH_2O \rightarrow nCO_2 + (3n + 1)H_2$. The primary reaction products are processed further in various ways, depending on the desired application of the hydrogen. Another important process for hydrogen production is the noncatalytic partial [oxidation](#) of hydrocarbons under elevated pressures: $C_nH_{2n+2} + (n/2)O_2 \rightarrow nCO + (n + 1)H_2$. This process requires a feed system for delivering precise rates of fuel and oxygen, burners of special design to give rapid mixing of the reactants, a refractory-lined reactor, and a [cooling system](#) to recover heat from the effluent gases. The latter process is exothermic (heat producing), in contrast to the endothermic (heat absorbing) steam–hydrocarbon process.

In a third process, called the pressure catalytic partial oxidation method, the two preceding processes are combined to maintain the required reaction temperature without external heating of the [catalyst](#) bed. Superheated steam and hydrocarbons are mixed, preheated, and blended with heated oxygen in a diffuser at the top of the catalytic reactor. The oxygen reacts with the hydrocarbons in a space above the catalyst. The reactants then pass through a bed of nickel catalyst in which the steam–hydrocarbon reactions proceed almost to [equilibrium](#).

Before 1940 most of the world production of hydrogen was made by processes based on coal or coke, the principal one being a [water–gas](#) reaction between steam and red-hot coke: $H_2O + C \rightarrow CO + H_2$. By 1970, however, relatively little hydrogen was being produced by such processes. For many years relatively small amounts of hydrogen had been produced by the [electrolysis](#) of aqueous solutions of [salt](#) or [sodium hydroxide](#), the electrode reaction being $H_2O + e^- \rightarrow \frac{1}{2}H_2 + OH^-$. The reaction between sulfuric or [hydrochloric acid](#) and an active [metal](#) like [zinc](#) is utilized to liberate hydrogen in the [laboratory](#), but such hydrogen usually contains trace quantities of volatile hydrides, such as [arsine](#) (AsH₃) and [phosphine](#)

(PH₃), produced by impurities in the metal. These volatile impurities may be removed by bubbling the mixture of gases through a solution of a strong oxidizing agent, such as [potassium permanganate](#).

A commercial method was developed for separating hydrogen from [carbon monoxide synthesis gas](#) by [diffusion](#). The gas flows under pressure through bundles of tiny hollow polyester fibres through whose walls the hydrogen passes.



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The largest single use of hydrogen in the world is in [ammonia](#) manufacture, which consumes about two-thirds of the world's hydrogen production. Ammonia is manufactured by the so-called [Haber-Bosch process](#), in which hydrogen and [nitrogen](#) react in the presence of a catalyst at pressures around 1,000 atmospheres and temperatures around 500° C: $N_2 + 3H_2 \rightarrow 2NH_3$. Large amounts of hydrogen are used in the preparation of [methanol](#) by the reaction $CO + 2H_2 \rightarrow CH_3OH$. This process is carried out in the presence of certain mixed [catalysts](#) containing zinc oxide and [chromium](#) oxide at temperatures between 300° and 375° C and at pressures between 275 and 350 atmospheres.

Another major application of hydrogen is in the catalytic [hydrogenation](#) of [organic compounds](#). Unsaturated vegetable and animal oils and fats are hydrogenated to make margarine and vegetable shortening. Hydrogen is used to reduce aldehydes, fatty acids, and esters to the corresponding alcohols. Aromatic [compounds](#) can be reduced to the corresponding saturated compounds, as in the conversion of [benzene](#) to cyclohexane and of phenol to cyclohexanol. Nitro compounds can be reduced easily to amines.

Hydrogen has been used as a primary rocket fuel for [combustion](#) with oxygen or [fluorine](#) and is favoured as a propellant for nuclear-powered rockets and space vehicles. Another increasing use of hydrogen is in the direct reduction of [iron](#) ores to [metallic iron](#) and in the reduction of the oxides of [tungsten](#) and [molybdenum](#) to the metals. A hydrogen (reducing) atmosphere is employed in the pouring of special castings, in the manufacture of [magnesium](#), in the annealing of metals, and for the cooling of large electric motors. Hydrogen was once used for inflating lighter-than-air vessels, such as dirigibles and balloons, but now [helium](#) is generally used because it is nonflammable. The [barrage](#) balloons used in England during [World War II](#), however, were filled with hydrogen. Liquid hydrogen is used in the laboratory to produce low temperatures.



Learn about new water molecule-splitting technology that separates hydrogen and oxygen

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Analysis

When atoms are excited, as in an electric discharge, they radiate light at discrete wavelengths that appear as lines in the [spectrum](#). Inasmuch as the wavelengths of atomic [spectral lines](#) are characteristic of the element, the atomic spectrum may be used for identifying the element. The simplest of all such spectra is that of hydrogen. [Johann Jakob Balmer](#), a Swiss mathematician and secondary school teacher, in 1885 discovered an equation for representing the wavelengths of hydrogen spectral lines, of which nine had been observed in the laboratory and of which five more were photographed in the spectrum of the [star](#) Sirius. The wavelengths, lambda (λ), in angstroms, were represented by the formula: $\lambda = 3645.6 [m^2/(m^2 - 4)]$, m taking the successive values 3, 4, 5, etc. It was not until 1913 that a theoretical basis for this [empirical](#) relation was given by the Danish physicist [Niels Bohr](#) in his theory of atomic radiation.

The spinning motion of the [proton](#) gives it [magnetic](#) properties and causes it to precess in an applied [magnetic field](#), much as a [spinning top](#) precesses in a gravitational field. The frequency at which a particular proton precesses is determined by its local electrical [environment](#) and by the strength of the applied magnetic field. When hydrogen compounds are irradiated with electromagnetic waves of a particular frequency, the phenomenon of [resonance](#) absorption occurs at magnetic field strengths that are different for each structurally (magnetically) distinguishable proton in the [compound](#). Thus, proton magnetic [resonance](#) makes it possible to distinguish the structural types of hydrogen atoms present; furthermore, the absorption peak intensities are proportional to the number of hydrogen atoms of each kind. The absorption peaks are often split, however, because of the magnetic interaction of the proton magnetic moments among themselves. Proton-magnetic-resonance measurements provide data for the investigation of chemical structure.

One method for determining the total hydrogen content of a substance is to oxidize the substance completely in a stream of pure oxygen, which reacts with the hydrogen to produce [water](#) vapour. The resulting vapours are passed through a powerful dehydrating agent, such as magnesium perchlorate, which absorbs the water. From the increase in weight of the absorption tube containing the desiccant, the amount of hydrogen oxidized can be calculated. Gaseous hydrogen or hydrogen compounds may be oxidized by passing them over hot [copper](#) oxide, and the resulting water can then be collected and weighed and the amount of hydrogen calculated; to measure the hydrogen gas itself, the water vapour from the oxidation may be reduced to hydrogen gas by passing it over hot [uranium](#) metal—the hydrogen then being measured in a simple device called a gas buret.

Strongly acidic hydrogen atoms (as in compounds such as HCl, HNO₃, H₂SO₄, etc.) can be determined in solution by adding measured amounts of a strong base, such as sodium hydroxide, NaOH, until the [acid](#) is neutralized, using an indicator to determine the end point. The net reaction is $H^+ + OH^- \rightarrow H_2O$. Weakly acidic hydrogen atoms (such as that attached to the oxygen in [ethanol](#), C₂H₅OH, and those attached to the nitrogen in acetamide, CH₃CONH₂) can be [converted](#) to [methane](#) (measured in a gas buret) by reaction with the methyl [Grignard reagent](#), CH₃MgI. Hydridic hydrogen atoms (as in NaBH₄, LiH, etc.) can be converted to molecular hydrogen (measured in a gas buret) by reaction with an aqueous acid.