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CHANGES THAT OCCUR IN THE BODY WHEN ACIDS ARE RAISED OR DROPPED FROM THE NORM

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Annotation

This article provides information on how acids change in the body in a state that rises or falls from the norm. This article is written in a comparative way in the case of using literature and internet resources.

Key words
acid, solution, sulfuric, nitric, hydrochloric.

Acid, any substance that in water solution tastes sour, changes the colour of certain indicators (e.g., reddens blue litmus paper), reacts with some metals (e.g., iron) to liberate hydrogen, reacts with bases to form salts, and promotes certain chemical reactions (acid catalysis). Examples of acids include the inorganic substances known as the mineral acids—sulfuric, nitric, hydrochloric, and phosphoric acids—and the organic compounds belonging to the carboxylic acid, sulfonic acid, and phenol groups. Such substances contain one or more hydrogen atoms that, in solution, are released as positively charged hydrogen ions (see Arrhenius theory). Broader definitions of an acid, to include substances that exhibit typical acidic behaviour as pure compounds or when dissolved in solvents other than water, are given by the Brønsted–Lowry theory and the Lewis theory. Examples of nonaqueous acids are sulfur trioxide, aluminum chloride, and boron trifluoride. Compare base.

Solution, in chemistry, a homogenous mixture of two or more substances in relative amounts that can be varied continuously up to what is called the limit of solubility. The term solution is commonly applied to the liquid state of matter, but solutions of gases and solids are possible. Air, for example, is a solution consisting chiefly of oxygen and nitrogen with trace amounts of several other gases, and brass is a solution composed of copper and zinc.

A brief treatment of solutions follows. For full treatment, see liquid: Solutions and solubilities.

Life processes depend in large part on solutions. Oxygen from the lungs goes into solution in the blood plasma, unites chemically with the hemoglobin in the red blood cells, and is released to the body tissues. The products of digestion also are carried in solution to the different parts of the body. The ability of liquids to dissolve other fluids or solids has many practical applications. Chemists take advantage of differences in solubility to separate and purify materials and to carry out chemical analysis. Most chemical reactions occur in solution and are influenced by the solubilities of the reagents. Materials for chemical manufacturing equipment are selected to resist the solvent action of their contents.

The liquid in a solution is customarily designated the solvent, and the substance added is called the solute. If both components are liquids, the distinction loses significance; the one present in smaller concentration is likely to be called the solute. The concentration of any component in a solution may be expressed in units of weight or volume or in moles. These may be mixed—e.g., moles per litre and moles per kilogram.

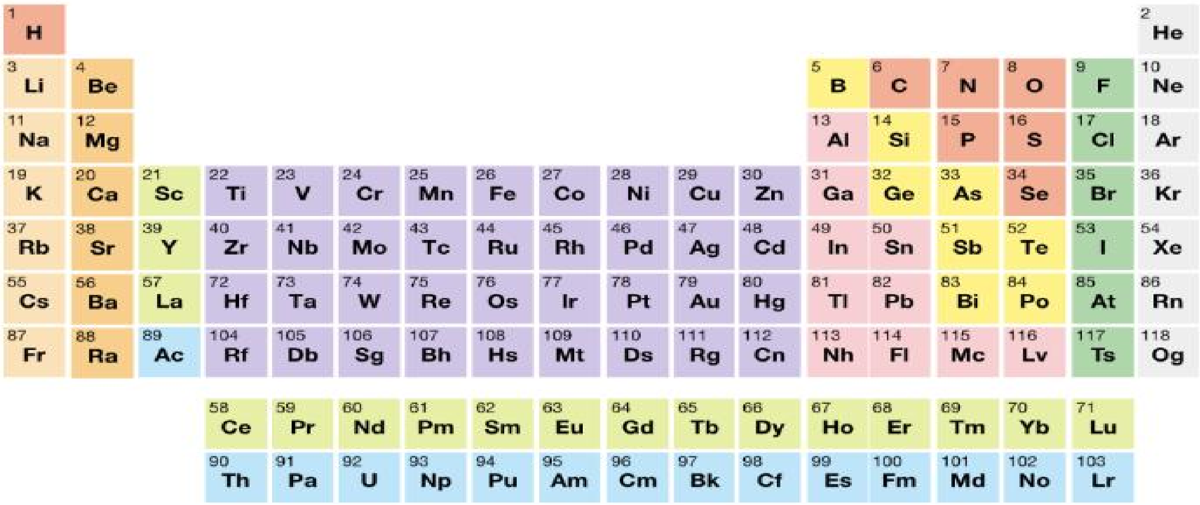
Crystals of some salts contain lattices of ions—i.e., atoms or groups of atoms with alternating positive and negative charges. When such a crystal is to be dissolved, the attraction of the oppositely charged ions, which are largely responsible for cohesion in the crystal, must be overcome by electric charges in the solvent. These may be provided by the ions of a fused salt or by electric dipoles in the molecules of the solvent. Such solvents include water, methyl alcohol, liquid ammonia, and hydrogen fluoride. The ions of the solute, surrounded by dipolar molecules of the solvent, are detached from each other and are free to migrate to charged electrodes. Such a solution can conduct electricity, and the solute is called an electrolyte.

The potential energy of attraction between simple, nonpolar molecules (nonelectrolytes) is of very short range; it decreases approximately as the seventh power of the distance between them. For electrolytes the energy of attraction and repulsion of charged ions drops only as the first power of the distance. AccordiIt is generally presumed that all gases are completely miscible (mutually soluble in all proportions), but this is true only at normal pressures. At high pressures, pairs of chemically dissimilar gases may very well exhibit only limited miscibility. Many different metals are miscible in the liquid state, occasionally forming recognizable compounds. Some are sufficiently alike to form solid solutions (see alloy).

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, H_2. The earliest known important chemical property of hydrogen is that it burns with oxygen to form water, H_2O; 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 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.



1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og
			58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
			90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

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.
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, H_2. 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.

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.

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