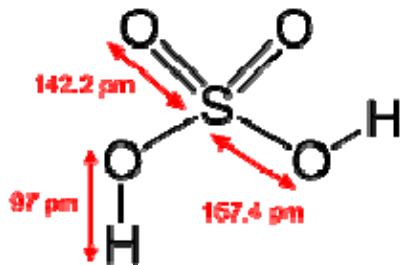




Sulfuric acid

Boiling Point: **338 °C (611 K)**

From Wikipedia, the free encyclopedia

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Sulfuric acid	
	
	
IUPAC name	Sulfuric acid
Other names	Oil of vitriol
Identifiers	
CAS number	[7664-93-9]
RTECS number	WS5600000

Properties	
Molecular formula	H ₂ SO ₄
Molar mass	98.078 g/mol
Appearance	clear, colorless, odorless liquid
Density	1.84 g cm ⁻³ , liquid
Melting point	10 °C, 283 K, 50 °F
Boiling point	290 °C, 563 K, 554 °F
Solubility in water	fully miscible (exothermic)
Viscosity	26.7 cP at 20°C
Hazards	
MSDS	External MSDS
EU classification	Corrosive (C)
NFPA 704	 0 3 2 COR
R-phrases	R35
S-phrases	(S1/2) , S26 , S30 , S45
Flash point	Non-flammable
Related compounds	

Related strong acids	Selenic acid Hydrochloric acid Nitric acid
Related compounds	Hydrogen sulfide Sulfurous acid Peroxymonosulfuric acid Sulfur trioxide Oleum
Supplementary data page	
Structure and properties	<i>n</i> , <i>ε</i> _r , etc.
Thermodynamic data	Phase behaviour Solid, liquid, gas
Spectral data	UV , IR , NMR , MS
Except where noted otherwise, data are given for materials in their standard state (at 25 °C, 100 kPa) Infobox references	

Sulfuric (or sulphuric) acid, H_2SO_4 , is a strong [mineral acid](#). It is soluble in [water](#) at all concentrations. It was once known as **oil of vitriol**, coined by the 8th-century [Muslim alchemist Jabir ibn Hayyan](#) (Geber) after his discovery of the chemical.^[1] Sulfuric acid has many applications, and is one of the top products of the [chemical industry](#). World production in 2001 was 165 million [tonnes](#), with an approximate value of US\$8 billion. Principal uses include [ore processing](#), [fertilizer manufacturing](#), [oil refining](#), [wastewater processing](#), and [chemical synthesis](#).

Many [proteins](#) are made of [sulfur](#)-containing amino [acids](#) (such as [cysteine](#) and [methionine](#)) which produce sulfuric acid when [metabolized](#) by the body.

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[hide]

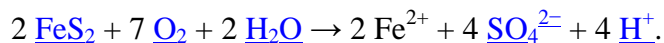
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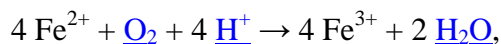
[[edit](#)] Occurrence

Pure (undiluted) sulfuric acid is not encountered on Earth, due to sulfuric acid's great affinity for water. Apart from that, sulfuric acid is a constituent of [acid rain](#), which is formed by atmospheric [oxidation](#) of [sulfur dioxide](#) in the presence of [water](#) - i.e., oxidation of [sulfurous acid](#). Sulfur dioxide is the main byproduct produced when sulfur-containing fuels such as coal or oil are burned.

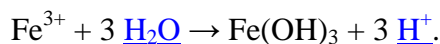
Sulfuric acid is formed naturally by the oxidation of sulfide minerals, such as iron sulfide. The resulting water can be highly acidic and is called [Acid Mine Drainage](#) (AMD). This acidic water is capable of dissolving metals present in sulfide ores, which results in brightly-colored, toxic streams. The oxidation of iron sulfide [pyrite](#) by molecular oxygen produces iron(II), or Fe²⁺:



The Fe²⁺ can be further oxidized to Fe³⁺, according to:



and the Fe³⁺ produced can be precipitated as the [hydroxide](#) or [hydrous oxide](#). The equation for the formation of the hydroxide is



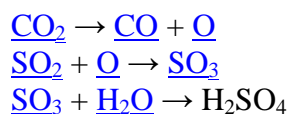
The iron(III) ion ("ferric iron", in casual nomenclature) can also oxidize pyrite. When iron(III) oxidation of pyrite occurs, the process can become rapid. pH values below zero have been measured in ARD produced by this process.

ARD can also produce sulfuric acid at a slower rate, so that the [Acid Neutralization Capacity](#) (ANC) of the aquifer can neutralize the produced acid. In such cases, the [Total](#)

Dissolved solids (TDS) concentration of the water can be increased from the dissolution of minerals from the acid-neutralization reaction with the minerals.

[edit] Extraterrestrial sulfuric acid

Sulfuric acid is produced in the upper atmosphere of [Venus](#) by the sun's [photochemical](#) action on [carbon dioxide](#), [sulfur dioxide](#), and [water](#) vapor. [Ultraviolet photons](#) of wavelengths less than 169 [nm](#) can [photodissociate carbon dioxide](#) into [carbon monoxide](#) and atomic [oxygen](#). Atomic oxygen is highly reactive. When it reacts with sulfur dioxide, a trace component of the Venerian atmosphere, the result is [sulfur trioxide](#), which can combine with water vapor, another trace component of Venus's atmosphere, to yield sulfuric acid.



In the upper, cooler portions of Venus's atmosphere, sulfuric acid exists as a liquid, and thick sulfuric acid [clouds](#) completely obscure the planet's surface when viewed from above. The main cloud layer extends from 45–70 [km](#) above the planet's surface, with thinner hazes extending as low as 30 and as high as 90 km above the surface.

Infrared spectra from [NASA's Galileo mission](#) show distinct absorptions on [Jupiter's moon Europa](#) that have been attributed to one or more sulfuric acid hydrates. The interpretation of the spectra is somewhat controversial. Some planetary scientists prefer to assign the spectral features to the sulfate ion, perhaps as part of one or more minerals on Europa's surface.^[2]

[edit] Manufacture

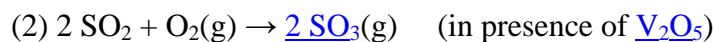
Main article: [Contact process](#)

Sulfuric acid is produced from [sulfur](#), [oxygen](#) and [water](#) via the [contact process](#).

In the first step, [sulfur](#) is burned to produce [sulfur dioxide](#).



This is then oxidized to [sulfur trioxide](#) using [oxygen](#) in the presence of a [vanadium\(V\) oxide catalyst](#).



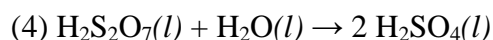
Finally the [sulfur trioxide](#) is treated with water (usually as 97-98% H_2SO_4 containing 2-3% water) to produce 98-99% sulfuric acid.



Note that directly dissolving SO_3 in water is not practical due to the highly [exothermic](#) nature of the [reaction](#), forming a corrosive mist instead of a liquid. Alternatively, SO_3 can be absorbed into H_2SO_4 to produce [oleum](#) ($\text{H}_2\text{S}_2\text{O}_7$), which may then be mixed with water to form sulfuric acid.



Oleum is reacted with water to form concentrated H_2SO_4 .



[\[edit\]](#) Physical properties

[\[edit\]](#) Forms of sulfuric acid

Although nearly 100% sulfuric acid can be made, this loses [\$\text{SO}_3\$](#) at the boiling point to produce 98.3% acid. The 98% grade is more stable in storage, and is the usual form of what is described as *concentrated sulfuric acid*. Other [concentrations](#) are used for different purposes. Some common concentrations are

- 10%, *dilute sulfuric acid* for laboratory use,
- 33.53%, *battery acid* (used in [lead-acid batteries](#)),
- 62.18%, *chamber* or *fertilizer acid*,
- 73.61%, *tower* or *Glover acid*,
- 97%, *concentrated acid*.

Different purities are also available. Technical grade H_2SO_4 is impure and often colored, but is suitable for making fertilizer. Pure grades such as [United States Pharmacopoeia](#) (USP) grade are used for making [pharmaceuticals](#) and [dyestuffs](#).

When high concentrations of $\text{SO}_3(\text{g})$ are added to sulfuric acid, $\text{H}_2\text{S}_2\text{O}_7$, called [pyrosulfuric acid](#), *fuming sulfuric acid* or [oleum](#) or, less commonly, *Nordhausen acid*, is formed. Concentrations of oleum are either expressed in terms of% SO_3 (called% oleum) or as% H_2SO_4 (the amount made if H_2O were added); common concentrations are 40% oleum (109% H_2SO_4) and 65% oleum (114.6% H_2SO_4). Pure $\text{H}_2\text{S}_2\text{O}_7$ is a solid with melting point 36°C .

[\[edit\]](#) Polarity and conductivity

[Anhydrous](#) H_2SO_4 is a very [polar](#) liquid, having a [dielectric constant](#) of around 100. It has a high electrical conductivity, caused by dissociation through [protonating](#) itself, a process known as [autoprotolysis](#).^[3]



The [equilibrium constant](#) for the autoprotolysis is^[3]

$$K_{\text{ap}}(25^\circ\text{C}) = [\text{H}_3\text{SO}_4^+][\text{HSO}_4^-] = 2.7 \times 10^{-4}.$$

The comparable equilibrium constant for [water](#), K_w is 10^{-14} , a factor of 10^{10} (10 billion) smaller.

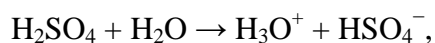
In spite of the viscosity of the acid, the effective [conductivities](#) of the H_3SO_4^+ and HSO_4^- ions are high due to an intra-molecular proton-switch mechanism (analogous to the [Grotthuss mechanism](#) in water), making sulfuric acid a good conductor. It is also an excellent solvent for many reactions.

The [equilibrium](#) is actually more complex than shown above; 100% H_2SO_4 contains the following species at equilibrium (figures shown as millimol per kg solvent): HSO_4^- (15.0), H_3SO_4^+ (11.3), H_3O^+ (8.0), HS_2O_7^- (4.4), [H₂S₂O₇](#) (3.6), H_2O (0.1).^[3]

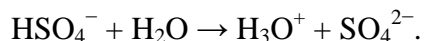
[\[edit\]](#) Chemical properties

[\[edit\]](#) Reaction with water

The [hydration reaction](#) of sulfuric acid is highly [exothermic](#). If [water](#) is added to the concentrated sulfuric acid, it can boil and spit dangerously. One should always add the acid to the water rather than the water to the acid. The necessity for this safety precaution is due to the relative densities of these two liquids. Water is less [dense](#) than sulfuric acid, meaning water will tend to float on top of this acid. The reaction is best thought of as forming [hydronium](#) ions, by



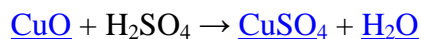
and then



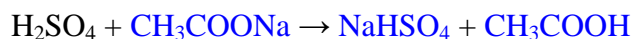
Because the hydration of sulfuric acid is [thermodynamically](#) favorable, sulfuric acid is an excellent dehydrating agent, and is used to prepare many dried fruits. The affinity of sulfuric acid for [water](#) is sufficiently strong that it will remove [hydrogen](#) and [oxygen](#) atoms from other compounds; for example, mixing [starch](#) $(\text{C}_6\text{H}_{12}\text{O}_6)_n$ and concentrated sulfuric acid will give elemental [carbon](#) and water which is absorbed by the sulfuric acid (which becomes slightly diluted): $(\text{C}_6\text{H}_{12}\text{O}_6)_n \rightarrow 6\text{C} + 6\text{H}_2\text{O}$. The effect of this can be seen when concentrated sulfuric acid is spilled on paper; the cellulose reacts to give a [burned](#) appearance, the [carbon](#) appears much as soot would in a fire. A more dramatic reaction occurs when sulfuric acid is added to a tablespoon of white [sugar](#); a rigid column of black, porous carbon will quickly emerge. The carbon will smell strongly of [caramel](#).

[\[edit\]](#) Other reactions

As an acid, sulfuric acid reacts with most [bases](#) to give the corresponding [sulfate](#). For example, [copper\(II\) sulfate](#). This blue salt of [copper](#), commonly used for [electroplating](#) and as a [fungicide](#), is prepared by the reaction of [copper\(II\) oxide](#) with sulfuric acid:

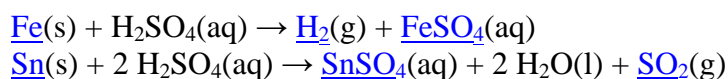


Sulfuric acid can also be used to displace weaker acids from their salts. Reaction with [sodium acetate](#), for example, displaces [acetic acid](#):

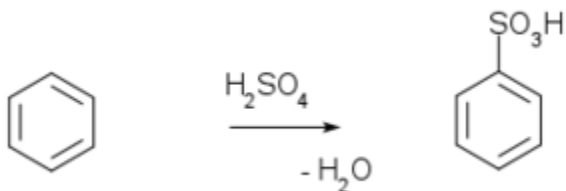


Similarly, reacting sulfuric acid with [potassium nitrate](#) can be used to produce [nitric acid](#) and a precipitate of [potassium bisulfate](#). When combined with [nitric acid](#), sulfuric acid acts both as an acid and a dehydrating agent, forming the [nitronium](#) ion NO_2^+ , which is important in [nitration](#) reactions involving [electrophilic aromatic substitution](#). This type of reaction, where protonation occurs on an [oxygen](#) atom, is important in many [organic chemistry](#) reactions, such as [Fischer esterification](#) and [dehydration of alcohols](#).

Sulfuric acid reacts with most metals via a single displacement reaction to produce [hydrogen](#) gas and the metal sulfate. Dilute H_2SO_4 attacks [iron](#), [aluminium](#), [zinc](#), [manganese](#), [magnesium](#) and [nickel](#), but reactions with [tin](#) and [copper](#) require the acid to be hot and concentrated. [Lead](#) and [tungsten](#), however, are resistant to sulfuric acid. The reaction with iron (shown) is typical for most of these metals, but the reaction with tin is unusual in that it produces [sulfur dioxide](#) rather than hydrogen.



Sulfuric acid undergoes [electrophilic aromatic substitution](#) with [aromatic compounds](#) to give the corresponding [sulfonic acids](#).^[4]

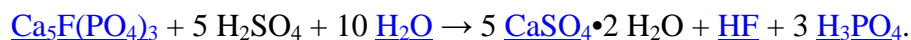


[\[edit\]](#) Uses



Sulfuric acid production in 2000

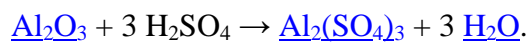
Sulfuric acid is a very important [commodity chemical](#), and indeed, a nation's sulfuric acid production is a good indicator of its industrial strength.^[5] The major use (60% of total production worldwide) for sulfuric acid is in the "wet method" for the production of [phosphoric acid](#), used for manufacture of [phosphate fertilizers](#) as well as [trisodium phosphate](#) for detergents. In this method, phosphate rock is used, and more than 100 million tonnes are processed annually. This raw material is shown below as [fluorapatite](#), though the exact composition may vary. This is treated with 93% sulfuric acid to produce [calcium sulfate](#), [hydrogen fluoride](#) (HF) and [phosphoric acid](#). The HF is removed as [hydrofluoric acid](#). The overall process can be represented as:



Sulfuric acid is used in large quantities by the [iron](#) and [steelmaking](#) industry to remove oxidation, [rust](#) and scale from rolled sheet and billets prior to sale to the [automobile](#) and white-goods industry. Used acid is often recycled using a Spent Acid Regeneration (SAR) plant. These plants combust spent acid with natural gas, refinery gas, fuel oil or other fuel sources. This combustion process produces gaseous sulfur dioxide (SO₂) and sulfur trioxide (SO₃) which are then used to manufacture "new" sulfuric acid. SAR plants are common additions to metal smelting plants, oil refineries, and other industries where sulfuric acid is consumed in bulk, as operating a SAR plant is much cheaper than the recurring costs of spent acid disposal and new acid purchases.

[Ammonium sulfate](#), an important nitrogen fertilizer, is most commonly produced as a byproduct from coking plants supplying the iron and steel making plants. Reacting the [ammonia](#) produced in the thermal decomposition of [coal](#) with waste sulfuric acid allows the ammonia to be crystallized out as a salt (often brown because of iron contamination) and sold into the agro-chemicals industry.

Another important use for sulfuric acid is for the manufacture of [aluminum sulfate](#), also known as paper maker's alum. This can react with small amounts of soap on [paper pulp](#) fibers to give gelatinous aluminum [carboxylates](#), which help to coagulate the pulp fibers into a hard paper surface. It is also used for making [aluminum hydroxide](#), which is used at [water treatment](#) plants to [filter](#) out impurities, as well as to improve the taste of the [water](#). Aluminum sulfate is made by reacting [bauxite](#) with sulfuric acid:



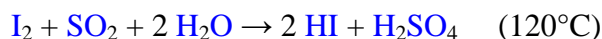
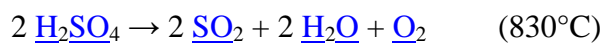
Sulfuric acid is used for a variety of other purposes in the chemical industry. For example, it is the usual acid catalyst for the conversion of [cyclohexanoneoxime](#) to

[caprolactam](#), used for making [nylon](#). It is used for making [hydrochloric acid](#) from [salt](#) via the [Mannheim process](#). Much H_2SO_4 is used in [petroleum](#) refining, for example as a catalyst for the reaction of [isobutane](#) with [isobutylene](#) to give [isooctane](#), a compound that raises the [octane rating](#) of [gasoline](#) (petrol). Sulfuric acid is also important in the manufacture of [dyestuffs](#) solutions and is the "acid" in lead-acid (car) batteries.

Sulfuric acid is also used as a general dehydrating agent in its concentrated form. See [Reaction with water](#).

[edit] Sulfur-iodine cycle

The [sulfur-iodine cycle](#) is a series of thermo-chemical processes used to obtain [hydrogen](#). It consists of three chemical reactions whose net reactant is [water](#) and whose net products are hydrogen and [oxygen](#).

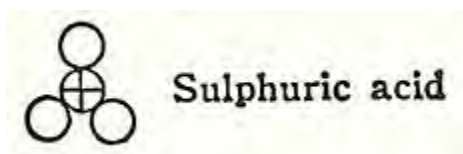



The sulfur and [iodine](#) compounds are recovered and reused, hence the consideration of the process as a cycle. This process is [endothermic](#) and must occur at high temperatures, so energy in the form of heat has to be supplied.

The sulfur-iodine cycle has been proposed as a way to supply hydrogen for a [hydrogen-based economy](#). It does not require [hydrocarbons](#) like current methods of [steam reforming](#).

The sulfur-iodine cycle is currently being researched as a feasible method of obtaining hydrogen, but the concentrated, corrosive acid at high temperatures poses currently insurmountable safety hazards if the process were built on large-scale.

[edit] History



 [John Dalton](#)'s 1808 sulfuric acid molecule shows a central sulfur atom bonded to three oxygen atoms.

The discovery of sulfuric acid is credited to the 8th century [Arabian chemist and alchemist, Jabir ibn Hayyan](#) (Geber). The acid was later studied by 9th century [Persian physician](#) and alchemist [Ibn Zakariya al-Razi](#) (Rhazes), who obtained the substance by [dry distillation](#) of minerals including [iron\(II\) sulfate](#) heptahydrate, $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$, and

[copper\(II\) sulfate](#) pentahydrate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. When heated, these compounds decompose to [iron\(II\) oxide](#) and [copper\(II\) oxide](#), respectively, giving off [water](#) and [sulfur trioxide](#), which combine to produce a dilute solution of sulfuric acid. This method was popularized in Europe through translations of Arabic and Persian treatises, as well as books by European alchemists, such as the 13th-century German [Albertus Magnus](#).

Sulfuric acid was known to medieval European alchemists as *oil of vitriol*, *spirit of vitriol*, or simply *vitriol*, among other names. The word vitriol derives from the Latin vitreus, 'glass', referring to the glassy appearance of the sulfate salts, which also carried the name vitriol. Salts called by this name included [copper\(II\) sulfate](#) (blue vitriol, or rarely [Roman vitriol](#)), [zinc sulfate](#) (white vitriol), [iron\(II\) sulfate](#) (green vitriol), [iron\(III\) sulfate](#) (vitriol of Mars), and [cobalt\(II\) sulfate](#) (red vitriol).

Vitriol was widely considered the most important [alchemical](#) substance, intended to be used as a [philosopher's stone](#). Highly purified vitriol was used as a medium for reacting other substances. This was largely because the acid does not react with [gold](#), production of which was often the final goal of alchemical processes. The importance of vitriol to alchemy is highlighted in the alchemical motto, *Visita Interiora Terrae Rectificando Invenies Occultum Lapidem* which is a [backronym](#) meaning ('Visit the interior of the earth and rectifying (i.e. purifying) you will find the hidden/secret stone'), found in *L'Azoth des Philosophes* by the 15th Century alchemist [Basilus Valentinus](#), .

In the 17th century, the German-Dutch chemist [Johann Glauber](#) prepared sulfuric acid by burning [sulfur](#) together with [saltpeter](#) ([potassium nitrate](#), KNO_3), in the presence of steam. As saltpeter decomposes, it oxidizes the sulfur to SO_3 , which combines with water to produce sulfuric acid. In 1736, Joshua Ward, a [London](#) pharmacist, used this method to begin the first large-scale production of sulfuric acid.

In 1746 in [Birmingham](#), [John Roebuck](#) adapted this method to produce sulfuric acid in [lead](#)-lined chambers, which were stronger, less expensive, and could be made larger than the previously used glass containers. This [lead chamber process](#) allowed the effective industrialization of sulfuric acid production. After several refinements, this method remained the standard for sulfuric acid production for almost two centuries.


Sulfuric acid created by John Roebuck's process only approached a 35–40% concentration. Later refinements to the lead-chamber process by French chemist [Joseph-Louis Gay-Lussac](#) and British chemist [John Glover](#) improved the yield to 78%. However, the manufacture of some [dyes](#) and other chemical processes require a more concentrated product. Throughout the 18th century, this could only be made by [dry distilling](#) minerals in a technique similar to the original [alchemical](#) processes. [Pyrite](#) ([iron disulfide](#), FeS_2) was heated in air to yield [iron \(II\) sulfate](#), FeSO_4 , which was oxidized by further heating in air to form [iron\(III\) sulfate](#), $\text{Fe}_2(\text{SO}_4)_3$, which, when heated to 480 °C, decomposed to [iron\(III\) oxide](#) and [sulfur trioxide](#), which could be passed through [water](#) to yield sulfuric acid in any concentration. However, the expense of this process prevented the large-scale use of concentrated sulfuric acid.

In 1831, British [vinegar](#) merchant Peregrine Phillips patented the [contact process](#), which was a far more economical process for producing sulfur trioxide and concentrated sulfuric acid. Today, nearly all of the world's sulfuric acid is produced using this method.

[\[edit\]](#) Safety

[\[edit\]](#) Laboratory hazards



 Drops of 98% sulfuric acid char a piece of tissue paper instantly

The corrosive properties of sulfuric acid are accentuated by its highly [exothermic reaction](#) with [water](#). Hence burns from sulfuric acid are potentially more serious than those of comparable strong acids (e.g. [hydrochloric acid](#)), as there is additional tissue damage due to dehydration and particularly due to the heat liberated by the reaction with water; i.e. secondary thermal damage. The danger is obviously greater with more concentrated preparations of sulfuric acid, but it should be remembered that even the normal laboratory "dilute" grade (approx. 1 M, 10%) will char paper by dehydration if left in contact for a sufficient time. Solutions equal to or stronger than 1.5 M should be labeled CORROSIVE, while solutions greater than 0.5 M but less than 1.5 M should be labeled IRRITANT. Fuming sulfuric acid (oleum) is not recommended for use in schools due to it being quite hazardous. The standard first aid treatment for acid spills on the skin is, as for other [corrosive](#) agents, irrigation with large quantities of water: Washing should be continued for at least ten to fifteen minutes in order to cool the tissue surrounding the acid burn and to prevent secondary damage. Contaminated clothing must be removed immediately and the underlying skin washed thoroughly.

Preparation of the diluted acid can also be dangerous due to the heat released in the dilution process. It is essential that the concentrated acid is added to water and not the other way round, to take advantage of the relatively high [heat capacity](#) of water. Addition of water to concentrated sulfuric acid leads at best to the dispersal of a sulfuric acid [aerosol](#), at worst to an [explosion](#). Preparation of solutions greater than 6 M (35%) in concentration is the most dangerous, as the heat produced can be sufficient to boil the diluted acid: efficient mechanical stirring and external cooling (e.g. an ice bath) are essential.

[\[edit\]](#) Industrial hazards

Although sulfuric acid is non-flammable, contact with metals in the event of a spillage can lead to the liberation of [hydrogen](#) gas. The dispersal of acid aerosols and gaseous [sulfur dioxide](#) is an additional hazard of fires involving sulfuric acid.

Sulfuric acid is not considered toxic besides its obvious corrosive hazard, and the main occupational risks are skin contact leading to burns (see above) and the inhalation of aerosols. Exposure to aerosols at high concentrations leads to immediate and severe irritation of the eyes, respiratory tract and mucous membranes: this ceases rapidly after exposure, although there is a risk of subsequent [pulmonary edema](#) if tissue damage has been more severe. At lower concentrations, the most commonly reported symptom of chronic exposure to sulfuric acid aerosols is erosion of the teeth, found in virtually all studies: indications of possible chronic damage to the [respiratory tract](#) are inconclusive as of 1997. In the [United States](#), the [permissible exposure limit](#) (PEL) for sulfuric acid is fixed at 1 mg/m³: limits in other countries are similar. Interestingly there have been reports of sulfuric acid ingestion leading to [vitamin B12 deficiency](#) with subacute combined degeneration. The spinal cord is most often affected in such cases, but the optic nerves may show [demyelination](#), loss of [axons](#) and [gliosis](#).

[\[edit\]](#) Legal restrictions

International commerce of sulfuric acid is controlled under the [United Nations Convention Against Illicit Traffic in Narcotic Drugs and Psychotropic Substances, 1988](#), which lists sulfuric acid under Table II of the convention as a chemical frequently used in the illicit manufacture of narcotic drugs or psychotropic substances.^[6]

In the [United States of America](#), sulfuric acid is included in [List II](#) of the [list of essential or precursor chemicals](#) established pursuant to the [Chemical Diversion and Trafficking Act](#). Accordingly, transactions of sulfuric acid—such as sales, transfers, exports from and imports to the United States—are subject to regulation and monitoring by the [Drug Enforcement Administration](#).^{[7][8][9]}

[\[edit\]](#) In fiction

In several films, cartoons and TV shows, especially [sci-fi](#) ones, sulfuric acid is normally depicted as a bubbling green steaming liquid, sometimes capable of dissolving almost anything in an instant. This is purely for visual appeal, since boiling green acid is more dangerous-looking than the actual clear and syrupy form it really is. The use of sulfuric acid as a weapon in crimes of assault, known as "[vitriol throwing](#)", has at times been sufficiently common (if sensational) to make its way into novels and short stories. Examples include *[The Adventure of the Illustrious Client](#)*, by [Arthur Conan Doyle](#), *[The Love of Long Ago](#)*, by [Guy de Maupassant](#) and *[Brighton Rock](#)* by [Graham Greene](#). A band, [My Vitriol](#), take their name from its use as a weapon in *Brighton Rock*. An episode of *[Saturday Night Live](#)* hosted by [Mel Gibson](#) included a parody Western sketch about "Sheriff Josh Acid," who carries a flask of acid instead of a six shooter. The [DC Comics](#) villain [Two Face](#) was disfigured as a result of a vitriol throw. This crime is also

mentioned in *Nineteen Eighty-Four* by [George Orwell](#); the protagonist Winston Smith agrees to throw vitriol into a child's face if that would be "the Brotherhood's" order, and Winston's enemy O'Brien later uses those barbaric words to undermine his logic. The novel *Veronika Decides to Die* by [Paulo Coelho](#) talks of a girl who has attempted to commit suicide and ends up with vitriol poisoning. The doctor/therapist in this novel also writes a thesis on curing vitriol poisoning. The substance was also used by a young gangster in Season 6B, Episode 5 of [The Sopranos](#) as a form of torture. It is also mentioned in Christian metal rock song "Acid Head" on [Tourniquet's](#) album *Vanishing Lessons* (1995).

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 - [^] [66 FR 52670—52675](#). 17 October 2001.
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 - [^] [21 USC, Chapter 13 \(Controlled Substances Act\)](#)
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[\[edit\]](#) External links

- [International Chemical Safety Card 0362](#)
- [NIOSH Pocket Guide to Chemical Hazards](#)
- [External Material Safety Data Sheet](#)
- [Sulfuric acid analysis - titration freeware](#)
- [- Sulfuric Acid density and pH-value at t=20°](#)

http://amazingrust.com/Experiments/how_to/Concentrating_H2SO4.html

Concentrating Sulfuric Acid (H_2SO_4)

Introduction



Sulfuric acid is an incredibly useful chemical and is readily available in the form of 'Battery Acid' used in Lead-Acid batteries (such as those found in automobiles). One can purchase Sulfuric / Battery acid at a very reasonable price at local automotive stores; this acid will usually have a concentration of about 30% to 35% H_2SO_4 by weight, the remainder being water.

For many applications, 30% H_2SO_4 will be more than concentrated enough to meet one's needs, however on occasion one might require a substantially higher concentration of Sulfuric acid. But purchasing concentrated Sulfuric acid can be much more costly and a greater hassle than buying battery acid, so a process which would enable one to concentrate Sulfuric acid oneself can sometimes be a desirable alternative. Fortunately, such a process exists and is relatively simple to perform.

One can concentrate Sulfuric acid by heating the liquid and boiling off water from the solution, leaving concentrated H_2SO_4 behind. By following this procedure, Sulfuric acid solutions may be concentrated to upwards of 98% H_2SO_4 by weight.

Safety

Safety is of paramount importance when concentrating Sulfuric Acid.

The concentrating process involves boiling Sulfuric acid to drive the water out of the solution. One should always handle Sulfuric acid with great care, but this is especially true when dealing with hot, concentrated, Sulfuric acid since it has the ability to act as an oxidizing acid. Needless to say, boiling Sulfuric acid is not something one would want to come into contact with due to the extremely corrosive nature of the acid, not to mention the injury incurred by a liquid boiling at several hundred degrees Celsius.

One absolutely must wear proper protective clothing.

Water (steam) is not the only vapor given off by the boiling acid. Along with water, Sulfur Dioxide and Trioxide (SO_2 and SO_3) gasses also evolve from the boiling liquid. At low acid concentrations, water is the predominant vapor given off by the boiling acid; significantly more water vapor evolves from the boiling liquid than SO_x gasses, thus allowing one to concentrate the acid through boiling. As the concentration rises, so does the amount of SO_x gasses given off by the acid. At higher concentrations the boiling acid begins to release, thick, toxic, SO_x fumes.

For this reason, one must perform the boiling outside in an area with plenty of ventilation and one should maintain a significant distance from the acid to avoid breathing in the harmful fumes.



One should only use undamaged, heat-resistant, borosilicate glass containers to boil the acid, such as Pyrex® (or, alternatively, Bomex) brand laboratory glassware. Do not use metal, plastic, or ordinary glass containers as they will not be able to withstand the corrosive nature of the acid or the high temperatures experienced during the boiling process.

A laboratory hot plate is the preferred heat source to boil the acid; an open-flame heat source is strongly discouraged as it places unnecessary stress on the glassware and may cause it to crack.

One would be wise to be prepared for the unfortunate (and hopefully unlikely) event of an accident occurring during the concentrating process, even after taking all the proper precautions. In case of a spill, having a bucket of water and an ample supply of Sodium Bicarbonate (baking soda) handy would be a very good idea in order to dilute and neutralize the acid.



Due to its highly corrosive nature and potential oxidizing properties, Sulfuric acid can be extremely dangerous when not handled responsibly. Always wear proper protective clothing when working around Sulfuric Acid (including, but not limited to, gloves and eye protection), and practice appropriate safety precautions.



Pre-Boiling

Pour the dilute acid to be concentrated into an appropriate boiling container (Pyrex® labware or equivalent) and begin to heat the liquid by gradually increasing the temperature setting on the heating apparatus.

Optionally, one may add boiling stones or other chemically inert objects to the container in order to better facilitate a controlled boiling process. The rough edges of the boiling stones assist in the formation of gas bubbles as molecules make

the transition into the vapor phase as the liquid begins to boil. Without this, it is possible for the liquid to become 'super-heated' as the liquid's temperature rises beyond its boiling point. The super-heated state can be easily disturbed and the liquid can suddenly begin to boil vigorously and can potentially boil over or splatter out of the container.

Boiling

When the acid begins to boil, water (steam) will be the most predominant vapor evolved from the solution. As the water leaves, the Sulfuric acid is left behind in a smaller and more concentrated solution.

As the acid becomes more concentrated, its boiling point will rise. Pure Sulfuric acid will boil at a temperature of 337 °C (610 K, 639 °F), whereas the more dilute the solution is, the closer its boiling point will be to that of water (100 °C).





Fuming

After a substantial portion of the water has boiled away, thick, white, fumes of SO_x gasses will begin to evolve from the boiling acid.

Depending on the initial concentration and volume of the dilute acid, it may require several hours of boiling in order to reach this stage.

The significantly increased concentration of the SO_x gasses in the fumes will be easy to differentiate from the ordinary water vapor evolving during the previous stage of the process. Compare the images of the boiling acid (above) to those of the fuming acid (below); the fumes from the “Boiling” stage are practically invisible whereas the fumes from the “Fuming” stage are thick and easily detectable.

The fumes containing the SO_x gasses will grow thicker as the concentration of the acid increases allowing one to use them as a crude method gauging the concentration of the acid.



Done

When the boiling acid's volume falls to about 20% - 25% of its original volume, one may stop boiling the acid and begin to let the liquid cool back down. Once the acid cools, carefully pour it into an appropriate storage container. Do not leave the cooled acid uncovered too long as Sulfuric acid is hygroscopic and will slowly absorb moisture from the air and dilute itself.

At this point, one should have obtained a Sulfuric acid solution which has a concentration somewhere around 95% H_2SO_4 . One can obtain a theoretical

maximum concentration of 98.3% H_2SO_4 by weight, but past this point the Sulfuric acid boils off with the water at a rate which does not allow the solution to become more concentrated.

If one wished to know more precisely the concentration of the recently prepared acid, one can measure the density of the Sulfuric acid and compare this measured value with that found in a chart containing the known H_2SO_4 concentration for a specified density. Additionally, one can perform a titration with an alkaline (basic) solution of a known concentration to calculate the H_2SO_4 concentration.

H_2SO_4 Density Table [from CRC Handbook of Chemistry and Physics, 81st edition, p. 8-81], showing molality (m, moles / kg), Molarity (c, mol / L), density (ρ , g / cm^3) for a given mass percent of Sulfuric acid solution.



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