

A collage background featuring various chemical structures and molecules. On the left, a yellow-green section shows the chemical formula $\text{H}_3\text{C}-\text{CH}_2-\text{CH}_2-\text{COOCH}_3$. In the center, a blue section shows NaCl and H_2O . On the right, a blue section shows OH and CO_2 . A magnifying glass is positioned over the H_2O molecule.

chemical COMPOUNDS



chemical
COMPOUNDS

chemical COMPOUNDS

NEIL SCHLAGER, JAYNE WEISBLATT, AND
DAVID E. NEWTON, *EDITORS*

Charles B. Montney, *Project Editor*

U·X·L

*An imprint of Thomson Gale,
a part of The Thomson Corporation*

THOMSON
★
GALE™



Chemical Compounds

Neil Schlager, Jayne Weisblatt, and David E. Newton, Editors

Project Editor

Charles B. Montney

Editorial

Luann Brennan, Kathleen J. Edgar,
Jennifer Greve, Madeline S. Harris, Melissa
Sue Hill, Debra M. Kirby, Kristine Krapp,
Elizabeth Manar, Kim McGrath, Paul
Lewon, Heather Price, Lemma Shomali

Indexing Services

Barbara Koch

Imaging and Multimedia

Randy Bassett, Michael Logusz

Product Design

Kate Scheible

Composition

Evi Seoud, Mary Beth Trimper

Manufacturing

Wendy Blurton, Dorothy Maki

© 2006 Thomson Gale, a part of the
Thomson Corporation.

Thomson and Star Logo are trademarks
and Gale and UXL are registered trade
marks used herein under license.

For more information, contact

Thomson Gale
27500 Drake Rd.
Farmington Hills, MI 48331 3535
Or you can visit our Internet site at
<http://www.gale.com>

ALL RIGHTS RESERVED

No part of this work covered by the
copyright hereon may be reproduced or
used in any form or by any means - gra-
phic, electronic, or mechanical, including
photocopying, recording, taping, Web
distribution, or information storage
retrieval systems - without the written
permission of the publisher.

For permission to use material from
this product, submit your request via
Web at <http://www.gale.edit.com/permissions>, or you may download our
Permissions Request form and submit
your request by fax or mail to:

Permissions

Thomson Gale
27500 Drake Rd.
Farmington Hills, MI 48331 3535
Permissions Hotline:
248 699 8006 or 800 877 4253, ext. 8006
Fax: 248 699 8074 or 800 762 4058

While every effort has been made to
ensure the reliability of the information
presented in this publication, Thomson
Gale does not guarantee the accuracy of
the data contained herein. Thomson
Gale accepts no payment for listing; and
inclusion in the publication of any orga-
nization, agency, institution, publica-
tion, service, or individual does not imply
endorsement of the editors or publisher.
Errors brought to the attention of the
publisher and verified to the satisfaction
of the publisher will be corrected in
future editions.

LIBRARY OF CONGRESS CATALOGING-IN-PUBLICATION DATA

Weisblatt, Jayne.
Chemical compounds / Jayne Weisblatt ; Charles B. Montney, project editor.
v. cm.
Includes bibliographical references and indexes.
Contents: v. 1. Acetaminophen through Dimethyl ketone -- v. 2. Ethyl acetate through
Polypropylene -- v. 3. Polysiloxane through Zinc oxide.
ISBN 1-4144-0150-7 (set : alk. paper) -- ISBN 1-4144-0451-4 (v. 1 : alk. paper) -- ISBN
1-4144-0452-2 (v. 2 : alk. paper) -- ISBN 1-4144-0453-0 (v. 3 : alk. paper)
1. Chemicals. 2. Organic compounds. 3. Inorganic compounds. I. Montney, Charles B.,
1962-. II. Title.
QD471.W45 2006
540--dc22

2005023636

This title is also available as an e book

ISBN 1 4144 0467 0

Contact your Thomson Gale sales representative for ordering information.

Printed in China

10 9 8 7 6 5 4 3 2 1

contents

volume I	<i>Reader's Guide</i>	xiii
<i>Timeline of the Development of</i>		
<i>Chemical Compounds</i>		xxi
<i>Words to Know</i>		xxxvii
1,3-Butadiene	1
2,2'-Dichlorodiethyl Sulfide	5
2-(4-Isobutylphenyl)propionic Acid	9
2,4,6-Trinitrotoluene	15
Acetaminophen	19
Acetic acid	23
Acetylene	27
Acetylsalicylic acid	31
Alpha-Tocopherol	37
Aluminum Fluoride	41
Aluminum Hydroxide	45
Aluminum Oxide	49
Aluminum Potassium Sulfate	53
Ammonia	57
Ammonium Chloride	63
Ammonium Hydroxide	69
Ammonium Nitrate	73

Contents

Ammonium Sulfate	77
Amoxicillin	81
Amyl Acetate	85
Amyl Nitrite	89
Ascorbic Acid	93
Benzene	99
Benzoic Acid	105
Beta-Carotene	109
Boric Acid	115
Butane	119
Butyl Acetate	125
Butyl Mercaptan	129
Butylated Hydroxyanisole and Butylated Hydroxytoluene	133
Caffeine	137
Calcium Carbonate	143
Calcium Hydroxide	147
Calcium Oxide	151
Calcium Phosphate	155
Calcium Silicate	161
Calcium Sulfate	165
Camphor	171
Carbon Dioxide	177
Carbon Monoxide	183
Carbon Tetrachloride	189
Cellulose	195
Cellulose Nitrate	201
Cellulose Xanthate	207
Chloroform	211
Chlorophyll	217
Cholesterol	223
Cinnamaldehyde	229
Citric Acid	233
Collagen	239

Copper(I) Oxide	243
Copper(II) Oxide	247
Copper(II) Sulfate	251
Cumene	255
Cyanoacrylate	259
Cyanocobalamin	265
Denatonium Benzoate	271
Dichlorodifluoromethane	277
Dichlorodiphenyltrichloroethane	283
Dimethyl Ketone	289
Ethyl Acetate	293
<i>Compounds by Formula</i>	lii
<i>Compounds by Element</i>	lix
<i>Compounds by Type</i>	lxv
<i>For Further Information</i>	lxix
<i>Index</i>	lxxix

volume 2 *Reader's Guide* xiii

<i>Timeline of the Development of Chemical Compounds</i>	xxi
<i>Words to Know</i>	xxxvii

Ethyl Alcohol	297
Ethylbenzene	303
Ethylene	307
Ethylene Glycol	313
Ethylene Oxide	317
Folic Acid	321
Formaldehyde	325
Fructose	329
Gamma-1,2,3,4,5,6-Hexachlorocyclohexane	333
Gelatin	337
Glucose	343

Contents

Glycerol	349
Hexane	353
Hydrogen Chloride	357
Hydrogen Peroxide	363
Iron(II) Oxide	367
Iron(III) Oxide	371
Isoamyl Acetate	377
Isoprene	381
Isopropyl Alcohol	387
Lactic Acid	391
Lactose	397
L-Aspartyl-L-Phenylalanine Methyl Ester	401
Luminol	407
Magnesium Chloride	411
Magnesium Hydroxide	415
Magnesium Oxide	419
Magnesium Silicate Hydroxide	423
Magnesium Sulfate	429
Menthol	435
Mercury(II) Sulfide	439
Methane	443
Methyl Alcohol	449
Methyl Mercaptan	455
Methyl- <i>t</i> -butyl Ether	459
Monosodium Glutamate	465
N,N-Diethyl-3-Methylbenzamide	469
Naphthalene	473
Naproxen	477
Niacin	483
Nicotine	487
Nitric Acid	493
Nitric Oxide	497
Nitrogen Dioxide	503
Nitroglycerin	507

Nitrous Oxide	513
Nylon 6 and Nylon 66	519
Oxalic Acid	525
Pectin	531
Penicillin	535
Perchlorates	541
Petrolatum	547
Petroleum	553
Phenol	559
Phosphoric Acid	565
Poly(Styrene-Butadiene-Styrene)	571
Polycarbonates	575
Polyethylene	579
Polymethyl Methacrylate	583
Polypropylene	587
Polysiloxane	591
<i>Compounds by Formula</i>	liii
<i>Compounds by Element</i>	lix
<i>Compounds by Type</i>	lxv
<i>For Further Information</i>	lxix
<i>Index</i>	lxxix

volume 3 *Reader's Guide* xiii

<i>Timeline of the Development of Chemical Compounds</i>	xxi
<i>Words to Know</i>	xxxvii

Polystyrene	597
Polytetrafluoroethylene	603
Polyurethane	609
Polyvinyl Chloride	615
Potassium Bicarbonate	621
Potassium Bisulfate	625

Contents

Potassium Bitartrate	629
Potassium Carbonate	633
Potassium Chloride	639
Potassium Fluoride	643
Potassium Hydroxide	647
Potassium Iodide	651
Potassium Nitrate	655
Potassium Sulfate	659
Propane	663
Propylene	669
Pyridoxine	673
Retinol	677
Riboflavin	683
Saccharin	689
Silicon Dioxide	695
Silver Iodide	701
Silver Nitrate	705
Silver(I) Oxide	711
Silver(I) Sulfide	715
Sodium Acetate	719
Sodium Bicarbonate	723
Sodium Carbonate	729
Sodium Chloride	735
Sodium Cyclamate	741
Sodium Fluoride	747
Sodium Hydroxide	753
Sodium Hypochlorite	759
Sodium Perborate	765
Sodium Phosphate	769
Sodium Polyacrylate	773
Sodium Silicate	779
Sodium Sulfite	785
Sodium Tetraborate	789
Sodium Thiosulfate	795

Stannous Fluoride	799
Styrene	803
Sucrose	807
Sucrose Polyester	813
Sulfur Dioxide	819
Sulfuric Acid	825
Tannic Acid	831
Testosterone	837
Theobromine	843
Thiamine	847
Toluene	853
Triclocarban	859
Triclosan	863
Urea	867
Vanillin	873
Water	879
Zinc Oxide	885
<i>Compounds by Formula</i>	iii
<i>Compounds by Element</i>	lix
<i>Compounds by Type</i>	lxv
<i>For Further Information</i>	lxix
<i>Index</i>	lxxix

reader's guide

Water; sugar; nylon; vitamin C. These substances are all very different from each other. But they all share one property in common: They are all chemical compounds. A chemical compound consists of two or more chemical elements, joined to each other by a force known as a chemical bond.

This book describes 180 chemical compounds, some familiar to almost everyone, and some less commonly known. Each description includes some basic chemical and physical information about the compound, such as its chemical formula, other names by which the compound is known, and the molecular weight, melting point, freezing point, and solubility of the compound. Here are some things to know about each of these properties:

Other Names: Many chemical compounds have more than one name. Compounds that have been known for many centuries often have common names that may still be used in industry, the arts, or some other field. For example, muriatic acid is a very old name for the compound now called hydrochloric acid. The name remains in common use today. Marine acid and spirit of salt are other ancient names for hydrochloric acid, but they are seldom used in the modern world. All compounds have systematic names, names based on a set of rules devised by the International Union of Pure and Applied Chemistry (IUPAC). For example, the systematic name for the poisonous gas whose common name is mustard gas is 2,2'-dichlorodiethyl sulfide. When chemists talk about chemical

compounds, they usually use only the official IUPAC name for a compound since that name leaves no doubt as to the substance about which they are talking. In some cases, a compound may have more than one official name, depending on the set of rules used in the naming process. For example, 1,1'-thiobis[2-chloroethane] is also an acceptable name for mustard gas. The “Other Names” section of each entry lists both the systematic (IUPAC) and common names for a compound.

Many compounds also have another kind of name, a brand name or trade name given to them by their manufacturers. For example, some trade names for the pain killer acetaminophen are PanadolTM, TylenolTM, AcetaTM, GenapapTM, TempraTM, and DepacinTM. The symbol next to each name means that the name is registered to the company that makes the compound. Trades names may be mentioned in the Overview or Uses sections of the entry for each compound.

Chemical Formula: A chemical formula is a set of symbols that tells the elements present in a compound and the relative numbers of each element. For example, the chemical formula for the compound carbon dioxide is CO₂. That formula tells that for every one carbon atom (C) in carbon dioxide there are two atoms of oxygen (O).

Chemists use different kinds of formulas to describe a compound. The simplest formula is a molecular formula. A molecular formula like CO₂ tells the kind and relative number of elements present in the compound. Another kind of formula is a structural formula. A structural formula provides one additional piece of information: The arrangement of elements in a compound. The structural formula for methanol (wood alcohol), for example, is CH₃OH. That formula shows that methanol consists of a carbon atom (C) to which are attached three hydrogen (H) atoms (CH₃). The carbon atom is also joined to an oxygen atom (O) which, in turn, is attached to a hydrogen atom (H).

Structural formulas can be written in a variety of ways. Another way to draw the structural formula for methanol, for example, is to show where individual bonds between atoms branch off other atoms in different directions. These structural formulas can be seen on the first page of nearly all entries in *Chemical Compounds*. In a third type of structural formula, the ball-and-stick formula, each element is

represented by a ball of some size, shape, and/or color. The chemical bond that holds them together is represented by sticks. This can be represented on paper in a drawing that simulates a three-dimensional model, by computer software, or actually in three dimensions from a kit with balls and sticks.

All three kinds of structural formulas are given for each compound described in this book. The only exception is some very large compounds known as polymers that contain many hundreds or thousands of atoms. In such cases, the formulas given shown only one small segment of the compound.

Compound Type: Millions of chemical compounds exist. To make the study of these compounds easier, chemists divide them into a number of categories. Nearly all compounds can be classified as either organic or inorganic. Organic compounds contain the element carbon; inorganic compounds do not. A few important exceptions to that rule exist, as indicated in the description of such compounds.

Both organic and inorganic compounds can be further divided into more limited categories, sometimes called families of compounds. Some families of organic compounds are the hydrocarbons (made of carbon and hydrogen only), alcohols (containing the -OH group), and carboxylic acids (containing the -COOH groups). Many interesting and important organic compounds belong to the polymer family. Polymers consist of very large molecules in which a single small unit (called the monomer) is repeated hundreds or thousands of times over. Some polymers are made from two or, rarely, three monomers joined to each other in long chains.

Most inorganic compounds can be classified into one of four major groups. Those groups are the acids (all of which contain at least one hydrogen (H) atom), bases (which all have a hydroxide (OH) group), oxides (which all have an oxygen (O)), and salts (which include almost everything else). A few organic and inorganic compounds described in this book do not easily fit into any of these families. They are classified simply as organic or inorganic.

Molecular Weight: The molecular weight of a compound is equal to the weight of all the elements of which it is made. The molecular weight of carbon dioxide (CO_2), for example, is equal to the atomic weight of carbon (12) plus two times

the atomic weight of oxygen ($2 \times 16 = 32$), or 44. Chemists have been studying atomic weights and molecular weights for a long time, and the molecular weights of most compounds are now known with a high degree of certainty. The molecular weights expressed in this book are taken from the *Handbook of Chemistry and Physics*, 86th edition, published in 2005. The Handbook is one of the oldest, most widely used, and most highly regarded reference books in chemistry.

Melting Point and Boiling Point: The melting point of a compound is the temperature at which it changes from a solid to a liquid. Its boiling point is the temperature at which it changes from a liquid to a gas. Most organic compounds have precise melting points and/or, sometimes, precise boiling points. This fact is used to identify organic compounds. Suppose a chemist finds that a certain unknown compound melts at exactly 16.5°C . Reference books show that only a small number of compounds melt at exactly that temperature (one of which is caprylic acid, responsible for the distinctive odor of some goats). This information helps the chemist identify the unknown compound.

Inorganic compounds usually do not have such precise melting points. In fact, they may melt over a range of temperatures (from 50°C to 55°C , for example) or sublime without melting. Sublimation is the process by which a substance changes from a solid to gas without going through the liquid phase. Other inorganic compounds decompose, or break apart, when heated and do not have a true melting point.

Researchers often find different melting points and boiling points for the same compound, depending on the reference book they use. The reason for this discrepancy is that many scientists have measured the melting points and boiling points of compounds. Those scientists do not always get the same result. So, it is difficult to know what the “true” or “most correct” value is for these properties. In this book, the melting points and boiling points stated are taken from the *Handbook of Chemistry and Physics*.

Some compounds, for a variety of reasons, have no specific melting or boiling point. The term “not applicable” is used to indicate this fact.

Solubility: The solubility of a compound is its tendency to dissolve in some (usually) liquid, such as water, alcohol, or

acetone. Solubility is an important property because most chemical reactions occur only when the reactants (the substances reacting with each other) are dissolved. The most common solvent for inorganic compounds is water. The most common solvents for organic compounds are the so-called organic solvents, which include alcohol, ether, acetone, and benzene. The solubility section in the entry for each compound lists the solvents in which it will dissolve well (listed as “soluble”), to a slight extent (“slightly soluble”), or not at all (“insoluble”).

Overview: The overview provides a general introduction to the compound, with a pronunciation of its name, a brief history of its discovery and/or use, and other general information.

How It Is Made: This section explains how the compound is extracted from the earth or from natural materials and/or how it is made synthetically (artificially). Some production methods are difficult to describe because they include reactants (beginning compounds) with difficult chemical names not familiar to most people with little or no background in chemistry. Readers with a special interest in the synthesis (artificial production) of these compounds should consult their local librarian or a chemistry teacher at a local high school or college for references that contain more information on the process in question. The For Further Information section may also contain this information.

Interesting Facts This section contains facts and tidbits of information about compounds that may not be essential to a chemist, an inventor, or some other scientific specialist, but may be of interest to the general reader.

Common Uses and Potential Hazards Chemical compounds are often of greatest interest because of the way they can be used in medicine, industry, or some other practical application. This section lists the most important uses of each compound described in the book.

All chemical compounds pose some risk to humans. One might think that water, sugar, and salt are the safest compounds in the world. But, of course, one can drown in water, become seriously overweight by eating too much sugar, and develop heart problems by using too much salt. The risk posed by a chemical compound really depends on a number of factors, one of the most important of which is the amount

of the compound to which one is exposed. The safest rule to follow in dealing with chemical compounds is that they are ALL dangerous under some circumstances. One should always avoid spilling any chemical compound on the skin, inhaling its fumes, or swallowing any of the compound. If an accident of this kind occurs, one should seek professional medical advice immediately. This book is not a substitute for prompt first aid properly applied.

Having said all that, some compounds do pose more serious health threats than others, and some individuals are at greater risks than others. Those special health risks are mentioned toward the end of the “Common Uses and Potential Hazards” section of each entry.

For Further Information As the name suggests, this section provides ideas for books, articles, and Internet sources that provide additional information on the chemical compound listed.

ADDED FEATURES

Chemical Compounds contains several features to help answer questions related to compounds, their properties, and their uses.

- The book contains three appendixes: a list by formula, list by element contained in compounds, and list by type of compound.
- Each entry contains up to two illustrations to show the relationship of the atoms in a compound to each other, one a black and white structural formula, and one a color ball-and-stick model of a molecule or portion of a molecule of the compound.
- A chronology and timeline in each volume locates significant dates in the development of chemical compounds with other historical events.
- “For Further Information,” a list of useful books, periodicals, and websites, provides links to further learning opportunities.
- The comprehensive index, which appears in each volume, quickly points readers to compounds, people, and events mentioned throughout *Chemical Compounds*.

ACKNOWLEDGMENTS

In compiling this reference, the editors have been fortunate in being able to rely upon the expertise and contributions of the following educators who served as advisors:

Ruth Mormon, Media Specialist, The Meadows School,
Las Vegas, Nevada

Cathy Chauvette, Sherwood Regional Library, Alexandria, Virginia

Jan Sarratt, John E. Ewing Middle School, Gaffney, South Carolina

Rachel Badanowski, Southfield High School, Southfield, Michigan

The editors would also like to thank the artists of Publishers Resource Group, under the lead of Farley Pedini, for their fast and accurate work and grace under pressure.

COMMENTS AND SUGGESTIONS

We welcome your comments on *Chemical Compounds*. Please write: Editors, *Chemical Compounds*, U•X•L, 27500 Drake Rd., Farmington Hills, MI 48331; call toll-free 1-800-877-4253; fax, 248-699-8097; or send e-mail via <http://www.gale.com>.

timeline of the development of chemical compounds

- c. 3000 BCE • Egyptians develop a method for converting gypsum to plaster of Paris, which is then used as mortar to join blocks in buildings.
- c. 2700 BCE • Chinese documents mention sodium chloride and the consumption of tea.
- c. 1550 BCE • The analgesic properties of willow tree bark, from which salicylic acid comes, are described in Egyptian scrolls.
- c. 1000 BCE • Ancient Egyptians use dried peppermint leaves.
- 800 BCE • Chinese and Arabic civilizations use borax for making glass and in jewelry work.

Timeline of the Development of Chemical Compounds

- 510 BCE** • Persian emperor Darius makes the first recorded reference to sugar when he refers to the sugar cane growing on the banks of the Indus River.
- 184 BCE** • Roman writer Cato the Elder describes a method of producing calcium oxide.
- c. 1st century CE** • Roman philosopher Pliny the Elder writes about a substance he calls hammoniacus sal, which appears to have been ammonium chloride.
- 1st century CE** • The first recipes calling for the use of pectin to make jams and jellies are recorded.
- c. 575 CE** • The cultivation of the coffee tree begins in Africa.
- 659** • Cinnamaldehyde is described in the famous Chinese medical text, the *Tang Materia Medica*.
- 8th century** • Arabian chemist Abu Musa Jabir ibn Hayyan, also known as Geber, writes about his work with several compounds, such as sodium chloride, sulfuric acid, nitric acid, citric acid, and acetic acid.
- 1242** • English natural philosopher Roger Bacon describes a method for making gunpowder.
- Late 1200s** • First mention of camphor by a Westerner occurs in the writings of Marco Polo.



Timeline of the Development of Chemical Compounds

- 1300s** • Potassium sulfate becomes known to alchemists.
- 1500s** • Spanish explorers bring vanilla to Europe from South and Central America, where it had already been used to flavor food.
- 1603** • Flemish chemist Jan Baptista van Helmont isolates a new gas produced during the combustion of wood, which is eventually called carbon dioxide.
- 1608** • Potash is one of the first chemicals to be exported by American colonists, with shipments leaving Jamestown, Virginia.
- 1610** • French alchemist Jean Béguin prepares acetone.
- 1620** • Flemish physician and alchemist Jan Baptista van Helmont first discovers nitric oxide.
- 1625** • German chemist Johann Rudolf Glauber is believed to have been the first to produce hydrogen chloride in a reasonably pure form. Later he is first to make ammonium nitrate artificially.
- 1695** • The term *Epsom salts* is introduced by British naturalist Nehemiah Grew, who names the compound after the spring waters near Epsom, England, from which it was often extracted.



1000

1200

1400

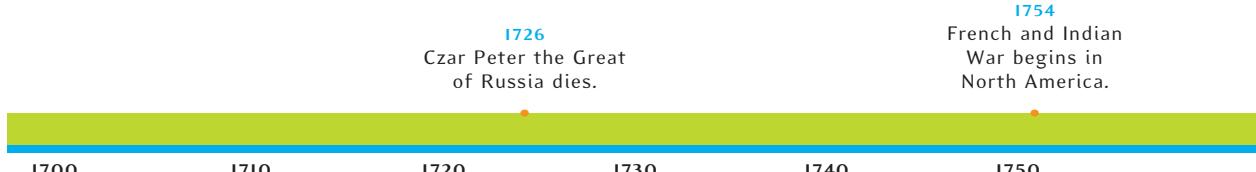
1500

1600

1700

Timeline of the Development of Chemical Compounds

- 1700** • German chemist Georg Ernst Stahl extracts acetic acid from vinegar by distillation.
- 1702** • German chemist Wilhelm Homberg is believed to be the first person to prepare boric acid in Europe.
- 1720s** • German chemist Johann Schulze makes discoveries that lead to using silver nitrate in printing and photography.
- 1746** • The first commercially successful method for making sulfuric acid is developed.
- 1747** • German chemist Andreas Sigismund Marggraf isolates a sweet substance from raisins that comes to be known as glucose.
- 1753** • James Lind reports that citrus fruits are the most effective means of preventing scurvy.
- 1769** • Oxalic acid is first isolated by German chemist Johann Christian Wieglob.
- 1770s** • British chemist Joseph Priestly does pioneering work with the compounds carbon dioxide, carbon monoxide, hydrogen chloride, and nitrous oxide, among others.
- 1770s** • Swedish chemist Karl Wilhelm Scheele discovers and works with phosphoric acid, glycerol, lactic acid, and potassium bitartrate.



Timeline of the Development of Chemical Compounds

1773 • French chemist Hilaire Marin Rouelle identifies urea as a component of urine.

Late 1700s • Commercial production of sodium bicarbonate as baking soda begins.

1776 • Carbon monoxide is first prepared synthetically by French chemist Joseph Marie François de Lassone, although he mistakenly identifies it as hydrogen.

1790 • The first patent ever issued in the United States is awarded to Samuel Hopkins for a new and better way of making pearl ash.

1794 • Ethylene is first prepared by a group of Dutch chemists.

Early 1800s • Silver iodide is first used in photography by French experimenter Louis Daguerre.

1817–1821 • French chemists Joseph Bienaimé Caventou and Pierre Joseph Pelletier successfully extract caffeine, quinine, strychnine, brucine, chinchonine, and chlorophyll from a variety of plants.

1817 • Irish pharmacist Sir James Murray uses magnesium hydroxide in water to treat stomach and other ailments. The compound is eventually called milk of magnesia.

1776
The U.S. Declaration of Independence is signed.

1776
French Revolution

1793
Cotton gin is invented by Eli Whitney.

1811 1812
Three severe earthquakes occur near New Madrid, Missouri.

1812
War of 1812 begins.

Timeline of the Development of Chemical Compounds

1818 • Hydrogen peroxide is discovered by French chemist Louis Jacques Thénard.

1819 • French naturalist Henri Braconnot discovers cellulose.

1825 • British chemist and physicist Michael Faraday discovers “bicarburet of hydrogen,” which is later called benzene.

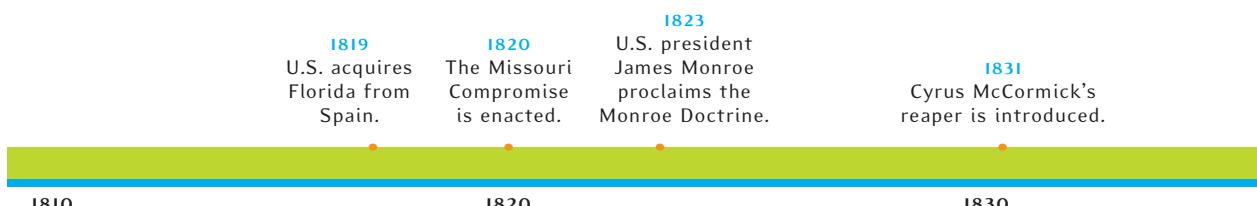
1830 • Peregrine Phillips, a British vinegar merchant from England, develops the contact process for making sulfuric acid. In the early 21st century it is still the most common way to make sulfuric acid.

1831 • Chloroform is discovered almost simultaneously by American, French, and German chemists. Its use as anesthetic is discovered in 1847.

1831 • Beta-carotene is first isolated by German chemist Heinrich Wilhelm Ferdinand Wackenroder.

1834 • Cellulose is first isolated and analyzed by French botanist Anselme Payen.

1835 • Polyvinyl chloride is first discovered accidentally by French physicist and chemist Henry Victor Regnault. PVC is rediscovered (again accidentally) in 1926.



Timeline of the Development of Chemical Compounds

- 1836** • British chemist Edmund Davy discovers acetylene.
- 1838** • French chemist Pierre Joseph Pelletier discovers toluene.
- 1839** • German-born French chemist Henri Victor Regnault first prepares carbon tetrachloride.
- 1839** • German druggist Eduard Simon discovers styrene in petroleum.
- 1845** • Swiss-German chemist Christian Friedrich Schönbein discovers cellulose nitrate.
- 1846** • Americans Austin Church and John Dwight form a company to make and sell sodium bicarbonate. The product will become known as Arm & Hammer® baking soda.
- Mid 1800s** • Hydrogen peroxide is first used commercially—primarily to bleach hats.
- 1850s** • Oil is first discovered in the United States in western Pennsylvania.

A horizontal timeline bar at the bottom of the page spans from approximately 1840 to 1860. It features four vertical orange dots corresponding to the years 1846, 1847, 1850, and 1858. Above each dot, a blue text box contains a historical event. The timeline is set against a background of alternating light blue and yellow horizontal stripes.

1846 Mexican American War begins.

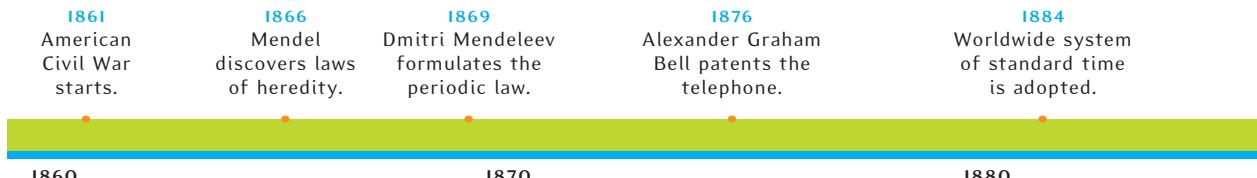
1847 Gold discovered in California.

1850 Levi Strauss manufactures his first pair of jeans.

1858 Lincoln debates Douglas in Illinois senate campaign.

Timeline of the Development of Chemical Compounds

- 1853** • French chemist Charles Frederick Gerhardt develops a method for reacting salicylic acid (the active ingredient in salicin) with acetic acid to make the first primitive form of aspirin.
- 1859** • Ethylene glycol and ethylene oxide are first prepared by French chemist Charles Adolphe Wurtz.
- 1860s** • Swedish chemist Alfred Nobel develops a process for manufacturing nitroglycerin on a large scale.
- 1863** • TNT is discovered by German chemist Joseph Wilbrand, although the compound is not recognized as an explosive until nearly 30 years later.
- 1865** • The use of carbolic acid as an antiseptic is first suggested by Sir Joseph Lister.
- 1865** • German botanist Julius von Sachs demonstrates that chlorophyll is responsible for photosynthetic reactions that take place within the cells of leaves.
- 1870** • American chemist Robert Augustus Chesebrough extracts and purifies petrolatum from petroleum and begins manufacturing it, eventually using the name VaselineTM.
- 1873** • German chemist Harmon Northrop Morse rediscovers and synthesizes acetaminophen. It had been discovered originally in 1852, but at the time it was ignored.



Timeline of the Development of Chemical Compounds

- 1879** • Saccharin, the first artificial sweetener discovered, is synthesized accidentally by Johns Hopkins researchers Constantine Fahlberg and Ira Remsen.
- 1879** • Riboflavin is first observed by British chemist Alexander Wynter Blyth.
- 1883** • Copper(I) oxide is the first substance found to have semiconducting properties.
- 1886** • American chemist Charles Martin Hall invents a method for making aluminum metal from aluminum oxide, which drastically cuts the price of aluminum.
- 1889** • French physiologist Charles E. Brown-Séquard performs early experiments on the effects of testosterone.
- 1890s** • Commercial production of perchlorates begins.
- 1890s–early 1900s** • British chemists Charles Frederick Cross, Edward John Bevan, and Clayton Beadle identify the compound now known as cellulose. They also develop rayon.
- Late 1890s** • Artificial methods of the production of pure vanillin are developed.
- 1901** • The effects of fluorides in preventing tooth decay are first observed.

1888
George Eastman
introduces the
Kodak camera.

1896
Henry Ford
assembles the
first motor car.

1901
The first Nobel
prizes are awarded.

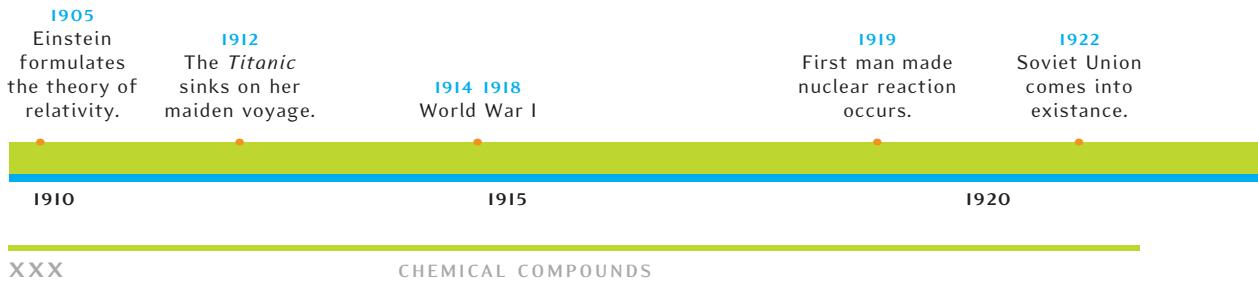
1890

1900

1910

Timeline of the Development of Chemical Compounds

- 1904** • German physicist Wilhelm Hallwachs discovers that a combination of copper metal and copper(I) oxide displays the photoelectric effect.
- 1910** • The first plant for the manufacture of rayon in the United States is built. By 1925 rayon becomes more popular than silk.
- 1910** • Japanese scientist Suzuki Umetaro discovers thiamine.
- 1912** • Nicotinic acid is first isolated by Polish-American biochemist Casimir Funk.
- 1914-1917** • Ethylene glycol is manufactured for use in World War I in explosives and as a solvent.
- 1914-1918** • A shortage of sugar during World War I leads to the reintroduction of saccharin to sweeten food. Saccharin production would boom again during World War II.
- Late 1910s** • Mustard gas is first used in war. Later it is found to be effective in treating cancer in experimental animals.
- 1922** • Vitamin E is discovered by two scientists at the University of California at Berkeley.



Timeline of the Development of Chemical Compounds

- 1928** • Polymethylmethacrylate (PMMA) is first synthesized in the laboratories of the German chemical firm Röhm and Haas.
- 1928** • Scottish bacteriologist Alexander Fleming accidentally discovers penicillin.
- 1930** • DuPont begins manufacturing dichlorodifluoromethane under the name Freon®.
- 1930** • A polymer based on styrene is produced by researchers at the German chemical firm I. G. Farben. Polystyrene comes to the United States in 1937.
- 1930s** • Properties and methods of synthesizing many vitamins, such as riboflavin, thiamine, niacin, ascorbic acid, pyridoxine, alpha-tocopherol, and retinol, are developed.
- Early 1930s** • SBS is first developed by German chemists Walter Bock and Eduard Tschunkur.
- 1933** • British chemists Reginald Gibson and Eric Fawcett accidentally re-discover polyethylene, which was first discovered in 1889.
- 1935** • Nylon is invented by Wallace Carothers. It is used in consumer products within three years.
- 1935** • German chemist Adolf Friedrich Johann Butenandt synthesizes testosterone.

1927
Charles Lindbergh
flies solo across the
Atlantic Ocean.

1929
The Great
Depression
begins

1925

1930

1935

Timeline of the Development of Chemical Compounds

1937 • German forensic scientist Walter Specht discovers that blood can act as the catalyst needed to produce chemiluminescence with luminol, a compound discovered in the late 1800s.

1937 • Plexiglas® (made from polymethylmethacrylate) is exhibited at the World's Trade Fair in Paris.

1937 • The basic process for making polyurethanes is first developed by German chemist Otto Bayer.

1937 • The cyclamate family of compounds is discovered by Michael Sveda, a graduate student at the University of Illinois.

1938 • Polytetrafluoroethylene is invented by Roy J. Plunkett by accident at DuPont's Jackson Laboratory.

1939 • Swiss chemist Paul Hermann Müller finds that DDT is very effective as an insecticide, which makes it useful in preventing infectious diseases such as malaria.

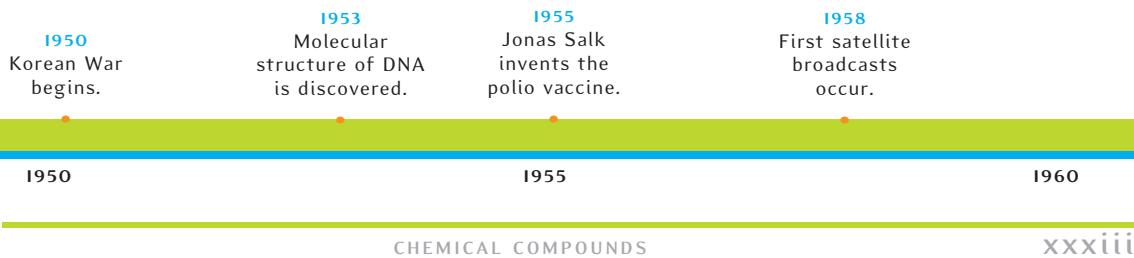
1939–1945 • During World War II, the U.S. military finds a number of uses for nylon, polyurethanes, polystyrene, perchlorates, and silica gel.

Early 1940s • Penicillin is first produced for human use and is valuable in saving the lives of soldiers wounded in World War II.



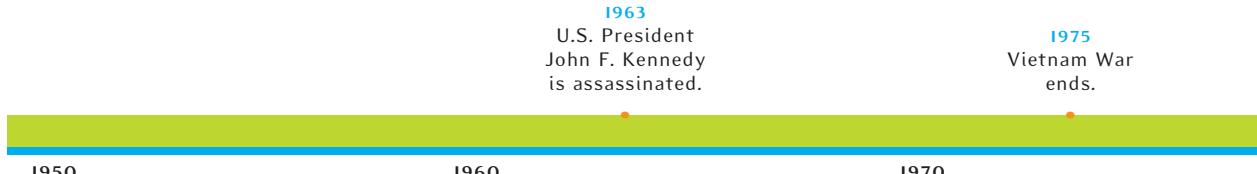
Timeline of the Development of Chemical Compounds

- 1940s** • A research chemist at the General Electric Company, E. G. Rochow, finds an efficient way of making organosiloxanes in large quantities.
- 1941** • Folic acid is isolated and identified by American researcher Henry K. Mitchell.
- 1941** • The first polyurethane adhesive, for the joining of rubber and glass, is made.
- 1942** • American researchers Harry Coover and Fred Joiner discover cyanoacrylate.
- 1946** • DEET is patented by the U.S. Army for use on military personnel working in insect-infested areas. It is made available to the public in 1957.
- 1947** • On April 16, an ammonium nitrate explosion in Texas City, Texas, becomes the worst industrial accident in U.S.
- 1950s** • Earliest reports surface about athletes using testosterone to enhance their sports performance.
- 1950s** • A stretchable material made of polyurethane, called spandex, is introduced.
- 1951** • Phillips Petroleum Company begins selling polypropylene under the trade name of Marlex®.
- 1953** • Polycarbonate, polyethylene, and synthetic rubber are developed.



Timeline of the Development of Chemical Compounds

- 1955** • Proctor & Gamble releases the first toothpaste containing stannous fluoride, Crest®.
- Mid 1950s** • Wham-O creates the hula-hoop—a ring of plastic that is made with low-grade polyethylene.
- 1956** • British chemist Dorothy Hodgkin determines the chemical structure of cyanocobalamin.
- 1958** • Scientist W. Barnes of the chemical firm T. & H. Smith in Edinburgh, Scotland, discovers denatonium benzoate.
- Early 1960s** • Ibuprofen is developed by researchers at the Boots Company, a British drug manufacturer.
- 1960s** • Triclosan becomes a common ingredient in soaps and other cleaning projects.
- 1960s** • MTBE is first synthesized by researchers at the Atlantic Richfield Corporation as an additive designed to increase the fuel efficiency of gasoline.
- 1962** • Amoxicillin is discovered by researchers at the Beecham pharmaceutical laboratories.
- 1965** • Aspartame is discovered accidentally by James M. Schlatter.



Timeline of the Development of Chemical Compounds

1980s • A ceramic form of copper(I) oxide is found to have superconducting properties at temperatures higher than previously known superconductors.

1980s • Polycarbonate bottles begin to replace the more cumbersome and breakable glass bottles.

1987 • Procter & Gamble seeks FDA approval of sucrose polyester. Ten years pass before the FDA grants that approval.

1994 • The U.S. Food and Drug Administration approves the sale of naproxen as an over-the-counter medication.

1995 • On April 19, American citizens Timothy McVeigh and Terry Nichols use a truckload of ammonium nitrate and other materials to blow up the Alfred P. Murrah Federal Building in Oklahoma City, Oklahoma.

Early 2000s • Some 350,000 propane-powered vehicles exist in the United States and about 4 million are used worldwide.

2004 • The leading chemical compound manufactured in the United States is sulfuric acid, with 37,515,000 metric tons (41,266,000 short tons) produced. Next is ethylene, with about 26.7 million metric tons (29.4 million short tons) produced.

1981
First personal
computers
become
available.

1989
The oil tanker
Exxon Valdez
sinks off
Alaska.

1991
Soviet Union
is dissolved.

2001
World Trade Center
in New York City
is destroyed.

1980

1990

2000

2006

words to know

A

ACETYL The organic group of acetic acid.

ADHESIVE A substance used to bond two surfaces together.

ALCHEMY An ancient field of study from which the modern science of chemistry evolved.

ALKALI A chemical base that can combine with an acid to produce a salt.

ALKALINE A substance that has a pH higher than 7.

ALKALOID An organic base that contains the element nitrogen.

ALKANE A type of hydrocarbon that has no double bonds because it contains the maximum possible number of hydrogen

ALKENE A kind of hydrocarbon with at least one double bond between carbons.

ALKYL GROUP A chemical group containing hydrogen and carbon atoms.

Words to Know

ALLOTROPE A form of an element that is different from its typical form, with a different chemical bond structure between atoms.

AMIDE An organic compound that includes the CON group bound to hydrogen.

AMINO ACID An organic compound that contains at least one carboxyl group (-COOH) and one amino group (-NH₂). They are the building blocks of which proteins are made.

ANALGESIC A substance that relieves pain.

ANHYDROUS Free from water and especially water that is chemically combined in a crystalline substance.

ANION A negatively charged ion.

ANODE The electrode in a battery in which electrons are lost (oxidized).

AROMATIC COMPOUND A compound whose chemical structure is based on that of benzene (C₆H₆).

B

BIODEGRADABLE Something that can be easily broken down by the action of bacteria.

BLOCK COPOLYMER A polymer composed of two or more different polymers, each of which clumps in blocks of identical molecules.

BORATE A salt that contains boron.

BRINE Salt water; water with a large amount of salt dissolved in it, such as seawater or water used to pickle vegetables.

BYPRODUCT A product that is made while making something else.

C

CARBOHYDRATES Organic compounds composed of carbon, oxygen, and hydrogen, which are used by the body as food.

CARBOXYL GROUPS Groups of atoms consisting of a carbon atom double bonded to an oxygen atom and single bonded to a hydroxyl (-OH) group (-COOH).

CARCINOGEN A substance that causes cancer in humans or other animals.

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

CATHODE The electrode in a battery through which electrons enter the fuel cell.

CATION A positively charged ion.

CAUSTIC Capable of burning or eating away, usually by the action of chemical reactions.

CENTRIFUGE A device that separates substances that have different densities by using centrifugal force.

CHELATE A chemical compound that is in the form of a ring. It usually contains one metal ion attached to a minimum of two nonmetal ions by coordinate bonds.

CHEMILUMINESCENCE Light produced by a chemical reaction.

CHIRAL A molecule with different left-handed and right-handed forms; not mirror symmetric.

CHLOROFLUOROCARBONS (CFCs) A family of chemicals made up of carbon, chlorine, and fluorine. CFCs were used as a refrigerant and propellant before they were banned for fear that they were destroying the ozone layer.

Words to Know

CHROMATOGRAPHY A process by which a mixture of substances passes through a column consisting of some material that causes the individual components in the mixture to separate from each other.

COAGULATE To make a liquid become a semisolid.

COENZYME A chemical compound that works along with an enzyme to increase the rate at which chemical reactions take place.

COMPOUND A substance formed of two or more elements that are chemically combined.

COPOLYMER A polymer made from more than one type of monomer.

COVALENT COMPOUND A compound in which the atoms are bonded to each other by sharing electrons.

CROSS-LINKED Polymer chains that are linked together to create a chemical bond.

CRYOGENICS The study of substances at very low temperatures, using substances such as liquefied hydrogen or liquefied helium.

D

DECOMPOSE To break a substance down into its most basic elements.

DENATURED To be made not fit for drinking.

DERIVATIVE Something gotten or received from another source.

DESICCANT Chemical agent that absorbs or adsorbs moisture.

DIATOMIC Composed of two atoms.

DISACCHARIDE A compound formed by the joining of two sugar molecules.

DISPERSANT A substance that keeps another substance from clumping together or becoming lumpy.

DISTILLATION A process of separating liquid by heating it and then condensing its vapor.

E

ELASTOMER A polymer known for its flexibility and elastic qualities; a type of rubber.

ELECTRODE A conductor through which an electric current flows.

ELECTROLYSIS A process in which an electric current is used to bring about chemical changes.

ELECTROLYTE A substance which, when dissolved in water, will conduct an electric current.

EMULSIFIER A substance that combines two other substances together that do not usually mix together and ensures they are spread evenly.

ESTER A compound formed by the reaction between an acid and an alcohol.

EXOTHERMIC Accompanied by the freeing of heat.

F

FAT An ester formed in the reaction between glycerol ($C_3H_5(OH)_3$) and a fatty acid, an organic acid with more than eight carbon atoms.

FERROUS Containing or made from iron.

FLOCCULANT A type of polymer, or large man-made particle, that is created by a repetitive chain of atoms.

FLUOROCARBON A chemical compound that contains carbon and fluorine, such as chlorofluorocarbon.

Words to Know

FOSSIL FUEL A fuel, such as petroleum, natural gas, or coal, formed from the compression of plant and animal matter underground millions of years ago.

FREE RADICAL An atom or group of atoms with a single unpaired electron that can damage healthy cells in the body.

G

G/MOL Grams per mole: a measure of molar mass that indicates the amount of the compound that is found in a mole of the compound. The molecular weight (also known as relative molecular mass) is the same number as the molar mass, but the g/mol unit is not used.

GLOBAL WARMING The increase in the average global temperature.

GLUCONEOGENESIS The production of glucose from non-carbohydrate sources, such as proteins and fats.

GLYCOGEN A carbohydrate that is stored in the liver and muscles, which breaks down into glucose.

GREENHOUSE EFFECT The increase of the average global temperature due to the trapping of heat in the atmosphere.

H

HARD WATER Water with a high mineral content that does not lather easily.

HELIX A spiral; a common shape for protein molecules.

HORMONE A chemical that delivers messages from one cell or organ to another.

HYDRATE A chemical compound formed when one or more molecules of water is added physically to the molecule of some other substance.

HYDROCARBON A chemical compound consisting of only carbon and hydrogen, such as fossil fuels.

HYDROGENATION A chemical reaction of a substance with molecular hydrogen, usually in the presence of a catalyst.

HYDROLYSIS The process by which a compound reacts with water to form two new compounds.

I

INCOMPLETE COMBUSTION Combustion that occurs in such a way that fuel is not completely oxidized.

INERT A substance that is chemically inactive.

INORGANIC Relating to or obtained from nonliving things.

ION An atom or molecule with an electrical charge, either positive or negative.

IONIC BOND A force that attracts and holds positive and negative ions together.

IONIC COMPOUND A compound that is composed of positive and negative ions, so that the total charge of the positive ions is balanced by the total charge of the negative ions.

ISOMER Two or more forms of a chemical compound with the same molecular formula, but different structural formulas and different chemical and physical properties.

ISOTOPE A form of an element with the usual number of protons in the nucleus but more or less than the usual number of electrons.

K

KREBS CYCLE A series of chemical reactions in the body that form part of the pathway by which cells break down carbohydrates, fats, and proteins for energy. Also called the citric acid cycle.

L

LATENT Lying hidden or undeveloped.

LEACH Passing a liquid through something else in order to dissolve minerals from it.

LEWIS ACID An acid that can accept two electrons and form a coordinate covalent bond.

LIPID An organic compound that is insoluble in water, but soluble in most organic solvents, such as alcohol, ether, and acetone.

M

MEGATON A unit of explosive force equal to one million metric tons of TNT.

METABOLISM All of the chemical reactions that occur in cells by which fats, carbohydrates, and other compounds are broken down to produce energy and the compounds needed to build new cells and tissues.

METALLURGY The science of working with metals and ores.

METHYLXANTHINE A family of chemicals including caffeine, theobromine, and theophylline, many of which are stimulants.

MINERALOGIST A scientist who studies minerals.

MISCIBLE Able to be mixed; especially applies to the mixing of one liquid with another.

MIXTURE A collection of two or more elements or compounds with no definite composition.

MONOMER A single molecule that can be strung together with like molecules to form a polymer.

MONOSACCHARIDE A simple sugar, made up of three to nine carbon atoms.

MORDANT A substance used in dyeing and printing that reacts chemically with both a dye and the material being dyed to help hold the dye permanently to the material.

N

NARCOTIC An addictive drug that relieves pain and causes drowsiness.

NEUROTRANSMITTER A chemical that relays signals along neurons.

NEUTRALIZE To make a substance neutral that is neither acidic or alkaline.

NITRATING AGENT A substance that turns other substances into nitrates, which are compounds containing NO_3 .

NSAID Non-steroidal anti-inflammatory drug, a drug that can stop pain and prevent inflammation and fever but that is not a steroid and does not have the same side effects as steroids.

O

ORGANIC Relating to or obtained from living things. In chemistry, refers to compounds made of carbon combined with other elements.

OXIDANT A substance that causes oxidation of a compound by removing electrodes from the compound. Also known as oxidizing agent.

OXIDATION STATE The sum of negative and positive charges, which indirectly indicates the number of electrons accepted or donated in the bond between elements.

Words to Know

OXIDIZE To combine a substance with oxygen, or to remove hydrogen from a molecule using oxygen, or to remove electrons from a molecule.

P

PARTICULATE MATTER Tiny particles of pollutants suspended in the air.

PETROCHEMICALS Chemical compounds that form in rocks, such as petroleum and coal.

PH (POTENTIAL HYDROGEN) The acidity or alkalinity of a substance based on its concentration of hydrogen ions.

PHOSPHATE A compound that is a salt of phosphoric acid, which consists of phosphorus, oxygen, sometimes hydrogen, and another element or ion.

PHOTOELECTRIC EFFECT The emission of electrons by a substance, especially metal, when light falls on its surface.

PHOTOSYNTHESIS The process by which green plants and some other organisms using the energy in sunlight to convert carbon dioxide and water into carbohydrates and oxygen.

PHOTOVOLTAIC EFFECT A type of photoelectric effect where light is converted to electrical voltage by a substance.

PLASTICIZER A substance added to plastics to make them stronger and more flexible.

POLYAMIDE A polymer, such as nylon, containing recurrent amide groups linking segments of the polymer chain.

POLYMER A substance composed of very large molecules built up by linking small molecules over and over again.

POLYSACCHARIDE A very large molecule made of many thousands of simple sugar molecules joined to each other in long, complex chains.

PRECIPITATE A solid material that settles out of a solution, often as the result of a chemical reaction.

PRECURSOR A compound that gives rise to some other compound in a series of reactions.

PROPRIETARY Manufactured, sold, or known only by the owner of the item's patent.

PROSTAGLANDINS A group of potent hormone-like substances that are produced in various tissues of the body. Prostaglandins help with a wide range of physiological functions, such as control of blood pressure, contraction of smooth muscles, and modulation of inflammation.

PROTEIN A large, complex compound made of long chains of amino acids. Proteins have a number of essential functions in living organisms.

Q

QUARRY An open pit, often big, that is used to obtain stone.

R

REAGENT A substance that is employed to react, measure, or detect with another substance.

REDUCTION A chemical reaction in which oxygen is removed from a substance or electrons are added to a substance.

REFRACTORY A material with a high melting point, resistant to melting, often used to line the interior of industrial furnaces.

RESIN A solid or semi-solid organic material that is used to make lacquers, adhesives, plastics, and many other clear substances.

S

SALT An ionic compound where the anion is derived from an acid.

SEMICONDUCTOR A material that has an electrical conductance between that of an insulator and a conductor. When charged with electricity or light, semiconductors change their state from nonconductive to conductive or vice versa.

SEQUESTERING AGENT A substance that binds to metals in water to prevent them from combining with other components in the water and forming compounds that could stain (sequestering agents are sometimes used in cleaning products). Also called a “chelating agent.”

SILICATE A salt in which the anion contains both oxygen and silicon.

SOLUBLE Capable of being dissolved in a liquid such as water.

SOLUTION A mixture whose properties are uniform throughout the mixture sample.

SOLVENT A substance that is able to dissolve one or more other substances.

SUBLIME To go from solid to gaseous form without passing through a liquid phase.

SUPERCONDUCTIVITY A state in which a material loses all electrical resistance: once established, an electrical current will flow forever.

SUPERCOOLED WATER Water that remains in liquid form, even though its temperature is below 0°C.

SUSPENSION A mixture of two substances that do not dissolve in each other.

SYNTHESIZE To produce a chemical by combining simpler chemicals.

T

TERATOGENIC Causing birth defects; comes from the Greek word “teratogenesis,” meaning “monster-making.”

THERMAL Involving the use of heat.

TOXIC Poisonous or acting like a poison.

TOXIN A poison, usually produced by microorganisms or by plants or animals.

TRAP A reservoir or area within Earth’s crust made of non-porous rock that can contain liquids or gases, such as water, petroleum, and natural gas.

U

ULTRAVIOLET LIGHT Light that is shorter in wavelength than visible light and can fade paint finishes, fabrics, and other exposed surfaces.

V

VASODILATOR A chemical that makes blood vessels widen, reducing blood pressure.

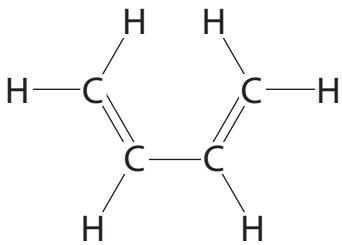
VISCOUS Having a syrupy quality causing a material to flow slowly.

VITRIFICATION The process by which something is changed into glass or a glassy substance, usually by heat.

VOLATILE Able to turn to vapor easily at a relatively low temperature.

W

WATER OF HYDRATION Water that has combined with a compound by some physical means.



OTHER NAMES:
Biethylene; bivinyl;
divinyl; erythrene;
vinylethylene

FORMULA:
 $\text{CH}_2=\text{CHCH}=\text{CH}_2$

ELEMENTS:
Carbon, hydrogen

COMPOUND TYPE:
Alkene (unsaturated
hydrocarbon)

STATE:
Gas

MOLECULAR WEIGHT:
54.09 g/mol

MELTING POINT:
−108.91°C
(−164.04°F)

BOILING POINT:
−4.41°C (−24.1°F)

SOLUBILITY:
Insoluble in water;
soluble in alcohol,
ether, and benzene

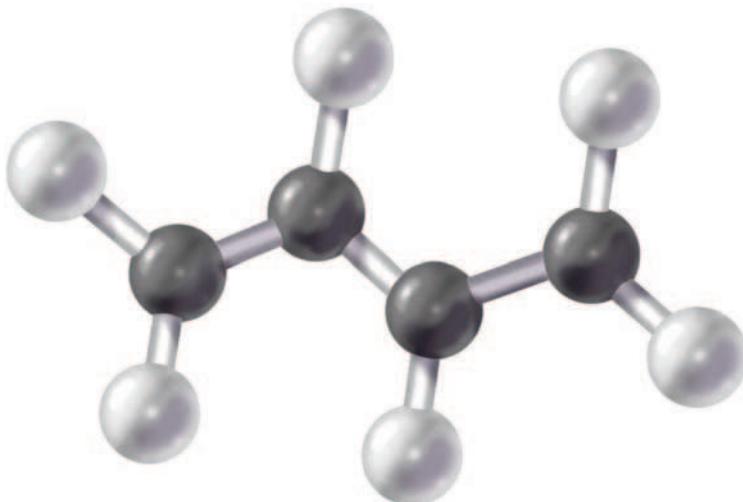
KEY FACTS

OVERVIEW

1,3-butadiene (one-three-byoo-tah-DYE-een) is a colorless gas with a mild, slightly sweet odor. It occurs naturally in petroleum, from which it is extracted at refineries. The compound was first discovered in petroleum in 1886 by the English chemist Henry E. Armstrong (1848–1937) and his colleague A. K. Miller (no dates available). This discovery apparently had no practical application until 1910 when the Russian chemist Sergei Lebedev (1874–1934) developed a method for polymerizing butadiene to make a rubber-like substance. Even then, Lebedev's invention, called polybutadiene, was primarily a laboratory curiosity. There was relatively little demand for rubber products that could not be met by the vast supplies of natural rubber from Southeast Asia.

That situation began to change in the 1920s and 1930s. The demand for rubber products in automobiles and other motor vehicles, especially tires, grew rapidly as car and truck production increased rapidly year after year. At the same

1,3 Butadiene. White atoms are hydrogen and black atoms are carbon. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP



time, manufacturers began to worry about losing their supply of natural rubber from traditional suppliers in Java, Malaysia, Sumatra, and other parts of Asia. As a result, chemists in Europe and the United States began to look for synthetic substitutes for natural rubber. Lebedev's discovery provided one solution to this problem, and butadiene-based synthetic rubber products were soon flowing from manufacturing plants.

HOW IT IS MADE

Butadiene is made by three processes in the United States. The first two of those processes begin with four-carbon compounds called butane ($\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$) and butene ($\text{CH}_2=\text{CHCH}_2\text{CH}_3$ and $\text{CH}_3\text{CH}=\text{CHCH}_3$). These compounds are treated in one of two ways so as to add one more double bond to the molecules, resulting in the formation of 1,3-butadiene ($\text{CH}_2=\text{CHCH}=\text{CH}_2$). The third and most common process is known as steam cracking. Steam cracking is a process by which large petroleum molecules are exposed to very hot steam, causing them to break apart into smaller molecules. Steam cracking is the method by which ethylene, the most important organic chemical produced in the United States, is made. 1,3-butadiene is a by-product of this process, and is separated from the major product, ethylene, after cracking is complete.

Interesting Facts

- In 2004, U.S. manufacturers produced 2.2 million metric tons (2.4 million short tons) of 1,3-butadiene, ranking it number 22 among all chemicals produced in the United States.

COMMON USES AND POTENTIAL HAZARDS

Virtually all of the 1,3-butadiene produced is used in the manufacture of polymers. Polybutadiene itself is too soft for most industrial rubber uses. So it is sometimes mixed with other polymers to make products with a variety of desirable qualities. Or it is polymerized with other monomers to make specialized copolymers. One of the most popular of the mixed polymers is polystyrene-polybutadiene. The second polymer in this mixture, polystyrene, is a polymer of the compound styrene ($C_6H_5CH=CH_2$). Some other uses of 1,3-butadiene include:

- Raw material in the manufacture of fungicides such as Captan and Captofol;
- Manufacture of latex adhesives;
- Production of domestic products, such as the backing on nylon carpet, appliance and electrical equipment, and luggage;
- Manufacture of industrial products, such as piping and conduits.

1,3-butadiene has a variety of harmful effects on humans, including irritation of the respiratory system, fatigue, drowsiness, headache, vertigo, loss of consciousness, respiratory paralysis, and death. The compound is also thought to be a carcinogen. Only people who work directly with the compound, however, are likely to experience such effects, and only then in unusual situations. The average person is likely to have no contact with the compound.

Words to Know

CARCINOGEN A substance that causes cancer in humans or other animals.

COPOLYMER A polymer made with two different monomers.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

POLYMERIZING The process by which polymers are made.

FOR FURTHER INFORMATION

“1,3 Butadiene.” National Safety Council.

http://www.nsc.org/library/chemical/13_Butad.htm (accessed September 14, 2005).

“1,3 Butadiene Health Effects.” Occupational Safety & Health Administration.

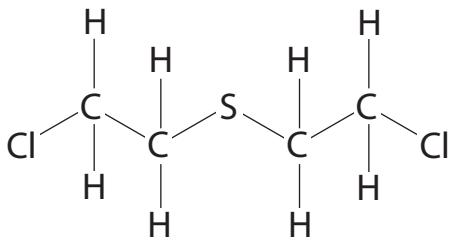
<http://www.osha.gov/SLTC/butadiene/healtheffects.html> (accessed September 14, 2005).

Johnson, Peter S. *Rubber Processing: An Introduction*. Cincinnati, OH: Hanser Gardner Publications, 2001.

“ToxFAQs for 1,3 Butadiene.” Agency for Toxic Substances and Disease Registry.

<http://www.atsdr.cdc.gov/tfacts28.html> (accessed September 14, 2005).

See Also Styrene



OTHER NAMES:

Mustard gas; see Overview for more names

FORMULA:
 $(\text{CH}_2\text{CH}_2\text{Cl})_2\text{S}$

ELEMENTS:
Carbon, hydrogen, chlorine, sulfur

COMPOUND TYPE:
Organic sulfide

STATE:
Liquid

MOLECULAR WEIGHT:
159.09 g/mol

MELTING POINT:
13 to 14°C (55 to 57°F)

BOILING POINT:
217°C (423°F)

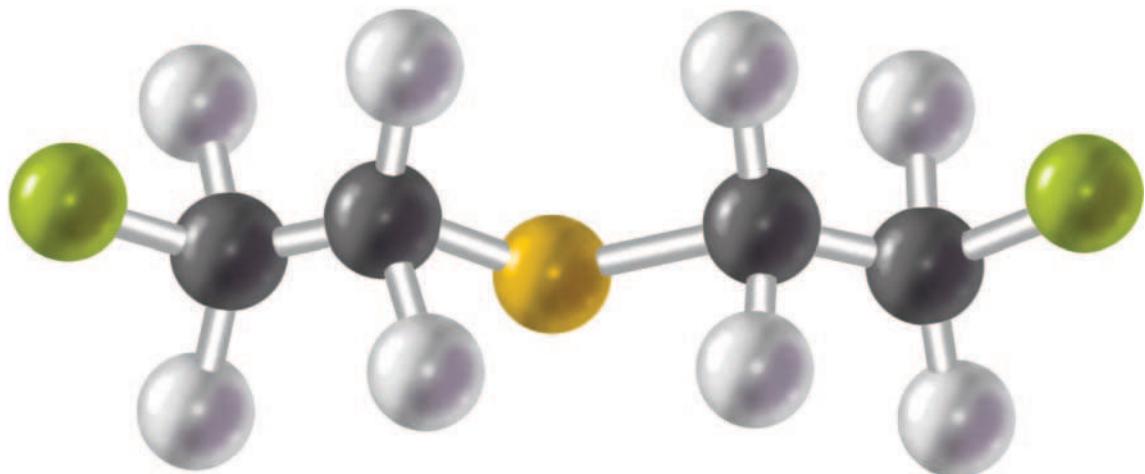
SOLUBILITY:
Insoluble in water;
soluble in alcohol,
ether, acetone,
benzene, and other
organic solvents;
soluble in fats

K E Y F A C T S

OVERVIEW

2,2'-dichlorodiethyl sulfide (two-two-prime-di-KLO-ro-di-ETH-el sul-fyd) is more commonly known as mustard gas. It is also known as bis-(2-chloroethyl) sulfide; sulfur mustard; yprite; and 1,1'-thiobis[2-chloroethane]. The compound occurs as a yellowish liquid that, in a pure form, has no odor. Small amounts of impurities give it the distinctive odor of mustard, from which it gets its common name. It may also smell like garlic or horseradish because of impurities.

Mustard gas was discovered by the English physicist Frederick Guthrie (1833–1886) in 1860. While working with the compound, Guthrie spilled mustard gas on his skin and found that it produced a painful red blister. That property has led to the primary use of mustard gas, as a chemical agent used in warfare. When sprayed on a person, the compound can blister the skin, burn the eyes, and irritate the lungs. In large doses, it can kill a person. Mustard gas was first used as a chemical weapon on July 12, 1917 by the German army in an attack on Canadian soldiers at Ypres,

***2,2' Dichlorodiethyl Sulfide.***

White atoms are hydrogen; black atoms are carbon; green atoms are chlorine; and yellow atom is sulfur.

PUBLISHERS RESOURCE GROUP

Belgium. An estimated 20,000 troops were killed in the attack. By the end of World War I, an estimated 120,000 British troops had died as a result of mustard gas attacks.

The damage caused by mustard gas was due at least in part because troops had no means of protecting themselves against the compound. It was able to penetrate virtually any type of protective clothing then available. The damage caused by mustard gas is so horrible that most nations have agreed not to use it in wars. One exception may have occurred during the Iraq-Iran war of 1980-1988, when the Iraqis sprayed Iranian troops with a gas very much like mustard gas in 1984. Later, in 1987 and 1988, Iraq's ruler Saddam Hussein also used mustard gas against Kurdish people living in northern Iraq.

HOW IT IS MADE

The usual method for making mustard gas in the United States is called the Levenstein process. In this process, ethylene gas ($\text{CH}_2=\text{CH}_2$) is bubbled through sulfur chloride (S_2Cl_2), a yellowish-red oil with a very strong odor. In Germany and other nations, the compound is made by treating the organic compound 2,2'-dihydroxyethyl sulfide with hydrochloric acid.

COMMON USES AND POTENTIAL HAZARDS

Virtually the only use for mustard gas is as a chemical agent in warfare. But, interestingly enough, the compound

Interesting Facts

- Another name for mustard gas, yprite, comes from the Belgian city Ypres (pronounced “EE-pr”), where the compound was first used as a military weapon.
- The use of mustard gas was outlawed by the Chemical Weapons Convention, adopted in 1993. The United States started destroying its stockpile of mustard gas even earlier, in 1985, and is now thought to have none of the chemical agent left.
- Mustard gas was not used during World War II. The compound was, nonetheless, responsible for the deaths of more than 100 U.S. men. On December 2, 1943, a German attack on the harbor of Bari, Italy, resulted in the sinking of about two dozen ships. Some of those ships were carrying mustard gas. When they sank, they released the gas into the surrounding water, where survivors inhaled and swallowed the poisonous compound.

has been used in one other very different way: in the treatment of cancer. The husband and wife medical team of E.B. and H.D. Krumbhaar discovered in 1919 that mustard gas was effective in treating cancer in experimental animals. (They called the compound by its then-popular name, yellow cross gas.) The compound was one of the first chemicals to be used in the treatment of cancer. Researchers continued to study the cancer-fighting properties of mustard gas for more than twenty years, but the compound has never been widely used for that purpose.

The effects of exposure to mustard gas vary over time. At first, a person who has inhaled the gas may become nauseated, experience itchy and sore eyes, and begin vomiting. His or her skin turns red and blisters may develop. Exposure to large amounts of the gas may result in permanent damage to the eyes and respiratory system resulting, in the worst cases, in coma and death. People exposed to mustard gas who survive may later develop mouth, lung, or stomach cancer.

2,2' DICHLORODIETHYL SULFIDE

FOR FURTHER INFORMATION

“Molecule of the Month: Mustard Gas.” Bristol University.

<http://www.chm.bris.ac.uk/motm/mustard/mustard.htm> (accessed on September 13, 2005).

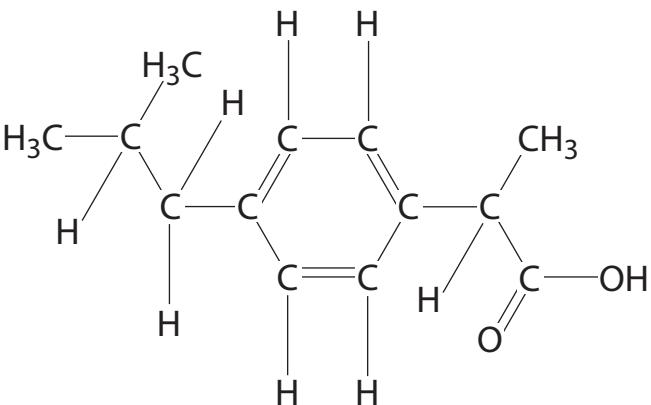
“Mustard Gas.” International Programme on Chemical Safety.

<http://www.inchem.org/documents/pims/chemical/mustardg.htm> (accessed on September 13, 2005).

Strobel, Warren P. “Saddam’s Lingering Atrocity.” U.S. News & World Report (November 27, 2000): 52.

“ToxFAQs for Sulfur Mustard.” Agency for Toxic Substances and Disease Registry.

<http://www.atsdr.cdc.gov/tfacts49.html> (accessed on September 13, 2005).



OTHER NAMES:

Ibuprofen; see Overview for more names

FORMULA:

C₁₃H₁₈O₂

ELEMENTS:

Carbon, hydrogen, oxygen

COMPOUND TYPE:

Organic acid

STATE:

Solid

MOLECULAR WEIGHT:

206.28 g/mol

MELTING POINT:

76°C (169°F)

BOILING POINT:

Not applicable;
decomposes above
melting point

SOLUBILITY:

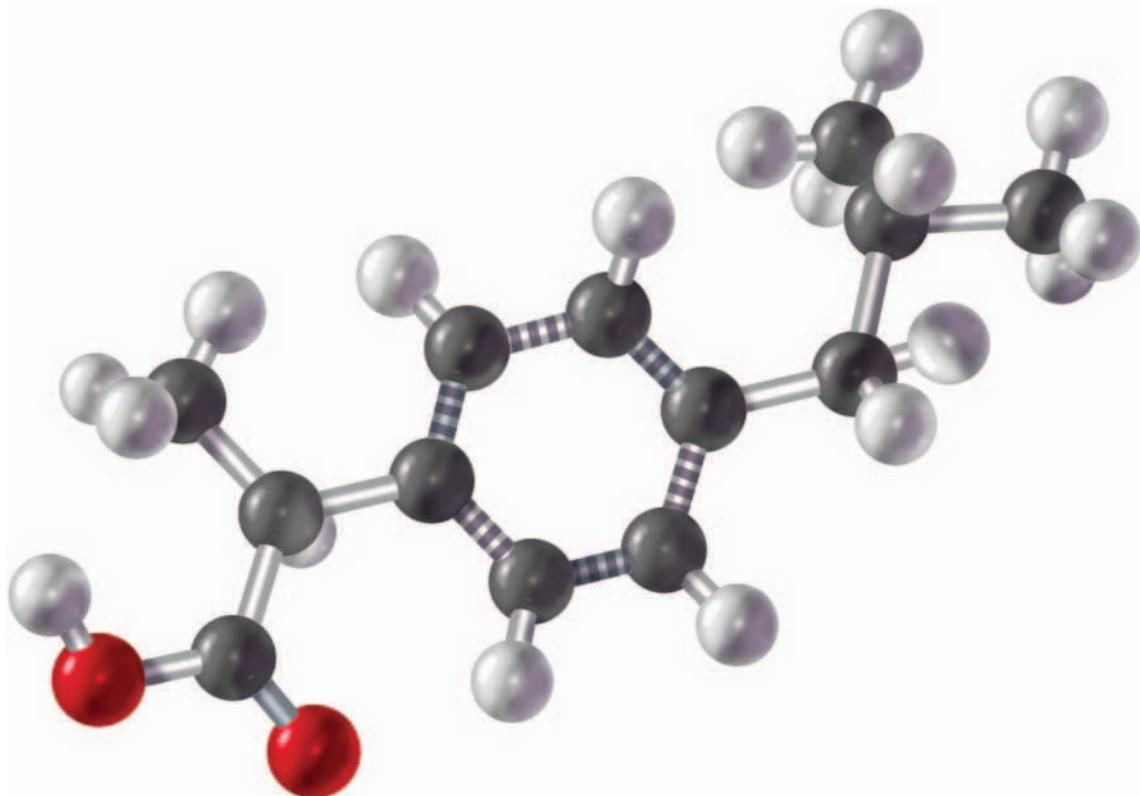
Slightly soluble in
water; soluble in most
organic solvents

KEY FACTS

OVERVIEW

2-(4-Isobutylphenyl)propionic acid (two four eye-so-BYOO-tuhl-PHEEN-uhl PRO-pi-on-ik AS-id) is a colorless crystalline solid widely used as an analgesic and anti-inflammatory medication. It is most commonly known as ibuprofen, but is also known as α -methyl-4-(2-methylpropyl)benzeneacetic acid and 4-isobutyl- α -methylphenylacetic acid. It is also sold under a variety of trade names, including Aches-n-Pain®, Advil®, Andran®, Antagil®, Antarene®, Excedrin IB®, Genpril®, Ibuprin®, Medipren®, Midol®, Motrin®, Nuprin®, PediaProfen®, Rufen®, and Vicoprofen®.

Ibuprofen was developed by researchers at the Boots Company, a British drug manufacturer, in the early 1960s. Those researchers had found that the anti-inflammatory property of aspirin was due to the presence of a carboxylic acid (-COOH) group in the compound. They searched for other carboxylic acids that might have similar properties and, after testing more than 600 compounds, discovered that ibuprofen met that criterion. It was twice as effective as



2 (4 isobutylphenyl) propionic acid. Red atoms are oxygen; white atoms are

hydrogen; and black atoms are carbon. Gray sticks indicate double bonds. Striped sticks indicate a benzene ring.

PUBLISHERS RESOURCE GROUP

acetylsalicylic acid (aspirin), the standard against which they were testing their products. They initially called their discovery "Brufen."

In 1974, Boots licensed the Upjohn Company in the United States to begin making a commercial version of Brufen. Upjohn called its product Motrin and marketed the compound primarily to people with arthritis. It soon became a popular drug with both doctors and patients, who found they were able to control the pain of arthritis with a non-addictive product that was easier on the stomach than aspirin.

Boots also entered the U.S. market several years later, selling ibuprofen under the trade name of Rufen. A number of other drug companies were eager to begin marketing their version of ibuprofen, and when the U.S. Food and Drug Administration approved ibuprofen for over-the-counter

Interesting Facts

Some evidence suggests that taking ibuprofen for two years or more may reduce the risk of developing Alzheimer's disease.

sales, they were ready. Whitehall Laboratories introduced Advil and Upjohn and Bristol-Meyers jointly released Nuprin. After Boots' patent ran out in 1985, several other drug manufacturers began producing their own versions of the popular pain-killing and anti-inflammatory product.

HOW IT IS MADE

The original method for making ibuprofen developed by Boots researchers begins with 2-methylpropylbenzene, a compound with a four-carbon methylpropyl group attached to a benzene ring. 2-methylpropylbenzene is a readily available product with a molecular structure quite similar to that of ibuprofen. Then, in a series of six steps, various groups of atoms are added, removed, or transposed to produce the 2-(4-isobutylphenyl)propionic acid molecule.

Other methods for preparing ibuprofen have also been developed. One of the most recent procedures has been invented by green chemists. Green chemists are chemists who look for methods of making chemical compounds that are less wasteful of raw materials, more sensitive to environmental concerns, and less hazardous to human health. The green chemistry method of making ibuprofen also begins with 2-methylpropylbenzene, but uses different catalysts and fewer steps to produce ibuprofen. Overall, the green method of making ibuprofen is twice as efficient (half as much waste products) as the original Boots method.

COMMON USES AND POTENTIAL HAZARDS

Ibuprofen is a nonsteroidal anti-inflammatory drug, or NSAID. The term nonsteroidal indicates that ibuprofen is not a steroid, a common class of drugs used to reduce

Words to Know

ANALGESIC a substance that relieves pain.

CATALYST a material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

GREEN CHEMISTRY the practice of making chemical compounds that are less wasteful of raw materials, more sensitive to environmental concerns, and less hazardous to human health.

inflammation and swelling. Scientists believe that ibuprofen works by inhibiting the enzyme cyclooxygenase (COX), which converts arachidonic acid to prostaglandins. Arachidonic acid is a naturally occurring fatty acid that is used to build a number of important biochemical compounds, including the prostaglandins. The prostaglandins are involved in the transmission of pain impulses throughout the nervous system. If the COX enzyme is prevented from functioning, arachidonic acid can not be converted into prostaglandins, and pain messages will not be transmitted.

Ibuprofen is especially effective in treating certain kinds of pain and inflammation, including those associated with menstrual cramps, various kinds of arthritis, headaches and migraines, pain from injuries and surgery, and discomfort associated with influenza and gout.

The major risk associated with use of ibuprofen is stomach irritation. The inventors of ibuprofen had hoped to find a substance that causes less stomach irritation than aspirin. Ibuprofen does seem to cause fewer gastrointestinal problems than does aspirin. Nevertheless, it can still cause severe problems in some patients, especially those who take large doses. Common complaints from the use of ibuprofen include nausea, stomach pain, and diarrhea. Patients who are especially sensitive or who take high doses of ibuprofen for long periods may experience damage to the stomach lining.

As with any drug, some patients should avoid taking ibuprofen entirely. These include individuals with heart or kidney problems, people who are taking so-called “blood thinners,” pregnant women during the last three months of

pregnancy, people with asthma, and people who are allergic to aspirin.

FOR FURTHER INFORMATION

"Ibuprofen." Chemical Land 21.

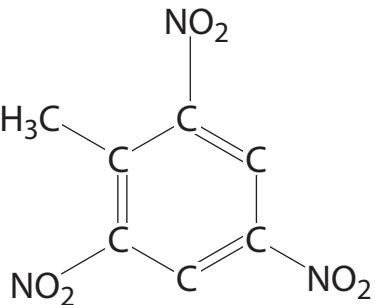
<http://www.chemicalland21.com/arokorhi/lifescience/phar/IBUPROFEN.htm> (accessed on November 3, 2005).

"Ibuprofen Oral." MedicineNet.com.

[http://www.medicinenet.com/ibuprofen oral/article.htm](http://www.medicinenet.com/ibuprofen_oral/article.htm)
(accessed on November 3, 2005).

Rainsford, K. D., ed. *Ibuprofen: A Critical Bibliographic Review*. Bethesda, MD: CCR Press, 1999.

See Also Acetylsalicylic Acid



OTHER NAMES:	Methyltrinitrobenzene; TNT
FORMULA:	C ₆ H ₂ (CH ₃)(NO ₂) ₃
ELEMENTS:	Carbon, hydrogen, nitrogen, oxygen
COMPOUND TYPE:	Substituted aromatic hydrocarbon (organic)
STATE:	Solid
MOLECULAR WEIGHT:	227.13 g/mol
MELTING POINT:	82°C (180°F)
BOILING POINT:	Decomposes explosively at 240°C (460°F)
SOLUBILITY:	Insoluble in water; soluble in alcohol, ether, acetone, and benzene

KEY FACTS

2,4,6-Trinitrotoluene

OVERVIEW

2,4,6-trinitrotoluene (two four six try-nye-troh-TOL-yoo-en), or TNT, is a yellow odorless solid that occurs in the form of crystalline needles. Its major use is in the manufacture of explosives. It may be used alone or in combination with other explosive chemicals. Although not as powerful as some other explosives, it has the advantage of being relatively insensitive to shock. Workers can handle the explosive without fear that it will suddenly explode if dropped or jarred. In fact, a blasting cap or detonator is needed to cause TNT to explode. Blasting caps and detonators are very sensitive explosives that can be attached to less sensitive explosives like TNT. Any small shock to the blasting cap or detonator will cause it to explode. That explosion, in turn, causes the less sensitive explosive, such as TNT, to blow up.

TNT was discovered by the German chemist Joseph Wilbrand (1811-1894) in 1863, although the compound was not recognized as an explosive until almost thirty years later. The compound was first produced commercially in Germany



2,4,6 Trinitrotoluene. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atoms are nitrogen. Gray sticks indicate double bonds. Striped sticks indicate a benzene ring.

PUBLISHERS RESOURCE GROUP

in 1901 and first used by Germany and its enemies during World War I. The explosive became so popular during the war that production of TNT could not keep up with demand by armies on both sides of the conflict.

HOW IT IS MADE

The method developed by Wilbrand for manufacturing TNT is simple, efficient, and inexpensive, and remains the

Interesting Facts

- During World War I (1914–1918), people who made explosives for the military (usually women and girls) were sometimes called “canary girls.” Their name came from the fact that their skin turned yellow when exposed to solid TNT. Women with red hair also found their hair turning an orangish or ginger color because of exposure to TNT.
- TNT has become the standard for explosive power in the world today. When someone talks about a “15 kiloton bomb” they are referring to a bomb with the explosive equivalent of 15 thousand (“kilo”) tons of TNT. Nuclear weapons commonly have explosive powers rated in the kiloton or megaton (million tons of TNT) range.

primary process by which TNT is produced today. Wilbrand made TNT by adding nitric acid (HNO_3) and sulfuric acid (H_2SO_4) to toluene ($\text{C}_6\text{H}_5\text{CH}_3$).

COMMON USES AND POTENTIAL HAZARDS

By far the most common use of TNT is in the manufacture of explosives. Until the discovery of nuclear energy in the 1940s, TNT was the most powerful explosive known to humans. Today, TNT is often combined with other explosives to make even more powerful bombs. Some examples include the following:

- Torpex, a mixture of TNT, wax and aluminum used in underwater explosives, such as those found in torpedoes. Torpex is about 50 percent more powerful than pure TNT;
- Pentolite, a combination of TNT and pentaerythritol tetranitrate (PETN) used primarily in blasting caps and detonators;
- Military dynamite, which contains the explosive RDX, TNT, motor oil, and cornstarch, is less powerful than pure dynamite, but much safer to handle;

Words to Know

CARCINOGEN A chemical that causes cancer in humans or other animals.

- Amatol, a combination of TNT and ammonium nitrate, is often substituted for TNT in ordnance (weapons);
- Minol, a variation of amatol that includes 20 percent aluminum to increase its explosive power; and
- Baratol, a mixture of TNT and barium nitrate that explodes more slowly than pure TNT.

In addition to its primary use in explosives, 2,4,6-trinitrotoluene is also used to produce a very attractive yellow tint in the manufacture of dyes and photographic chemicals.

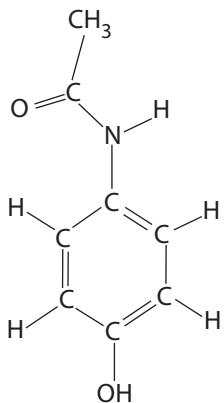
2,4,6-trinitrotoluene is a hazardous chemical that can produce both short- and long-term health effects, such as respiratory problems, skin irritation, anemia, and lung damage. Some research shows the compound may be carcinogenic in laboratory animals, but not in humans. Since most people are never exposed to TNT, these hazards are primarily of concern to men and women engaged in the manufacture or use of TNT.

FOR FURTHER INFORMATION

Cooper, P. W., and S. R. Kurowski. *Introduction to the Technology of Explosives*. New York: Wiley VCH, 1997.

“Public Health Statement for 2,4,6 Trinitrotoluene.” Agency for Toxic Substances and Disease Registry.

<http://www.atsdr.cdc.gov/toxprofiles/phs81.html> (accessed September 15, 2005).

**OTHER NAMES:**

See Overview.

FORMULA:

$\text{CH}_3\text{CONHC}_6\text{H}_4\text{OH}$

ELEMENTS:

Carbon, hydrogen, oxygen, nitrogen

COMPOUND TYPE:

Aromatic amide
(organic)

STATE:

Solid

MOLECULAR WEIGHT:

151.16 g/mol

MELTING POINT:

169°C-172°C
(336°F-342°F)

BOILING POINT:

Not applicable
(decomposes)

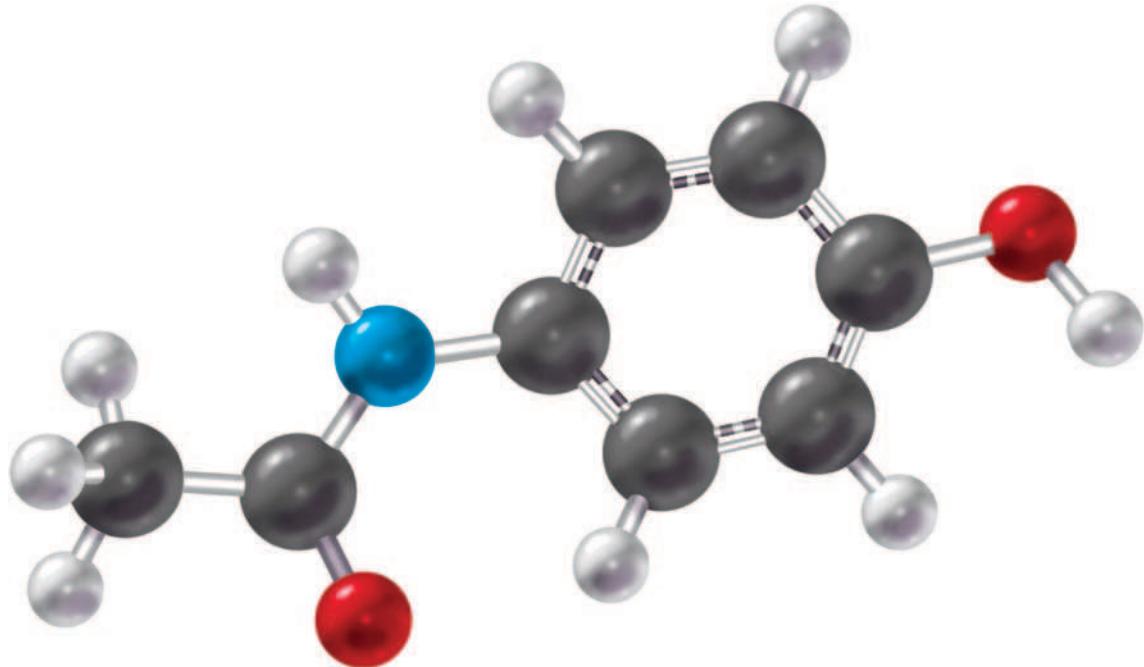
SOLUBILITY:

Insoluble in cold water; more soluble in warm water; soluble in alcohol, acetone, and other organic solvents

K E Y F A C T S**OVERVIEW**

Acetaminophen (uh-SEE-toe-Min-oh-fen) is one of the most commonly used analgesics in the world. It is also known as p-Acetylaminophenol; p-acetamidophenol; N-acetyl-p-aminophenol; p-hydroxyacetanilide; and paracetamol. It is an ingredient in more than 100 commercial products including Abensanil®, Acamol®, Acetagesic®, Alpinyl®, Alvedon®, Anaflon®, Anelix®, Anhiba®, Calpol®, Datril®, Dirox®, Doliprane®, Dymadon®, Enelfa®, Eneril®, Exadol®, Febrilix®, Febrolin®, Fendom®, Finimal®, Hedex®, Homoolan®, Lonarid®, Multin®, Panadol®, Phendon®, Tylenol®, Valdol®, and Valgesic®. Acetaminophen is often combined with other ingredients in medications such as cold and flu products, cough suppressants, and allergy medications in order to treat more than one symptom at a time.

Acetaminophen was first discovered in 1852 by the French chemist Charles Frederic Gerhardt, who also discovered aspirin a year later. His results were largely ignored, however, until a German chemist, Harmon Northrop Morse



Acetaminophen. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atom is nitrogen. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

(1848–1920) re-discovered and synthesized the compound in 1873. Even then, the medicinal qualities of the compound were not appreciated for another twenty years, and the compound was not even prescribed or used by doctors until 1949. Since that time, however, it has become extremely popular, exceeded only by aspirin as the most popular analgesic sold in the world.

HOW IT IS MADE

Acetaminophen is made in a straight-forward reaction between p-aminophenol [$C_6H_4(OH)(NH_2)$] and acetic anhydride [$(CH_3CO)_2O$]. As its formula shows, the p-aminophenol provides the basic structure from which acetaminophen is formed, with the acetic anhydride adding the $-CH_3CO$ group of atoms needed to complete the compound.

COMMON USES AND POTENTIAL HAZARDS

In addition to its use as a pain reliever, acetaminophen has a number of industrial uses, including:

Interesting Facts

In September 1982, seven individuals in the Chicago area died after taking a form of Tylenol® called Extra Strength Tylenol®. Officials later found that someone had added cyanide to the Tylenol® bottles. It was the

poison and not the pain killer that had caused the deaths. The crime was never solved, but it led to the introduction of so called tamper proof packaging for virtually all medicines.

- Stabilizer for hydrogen peroxide solutions, to prevent the peroxide from breaking down too quickly;
- Raw material in the production of other pharmaceutical compounds;
- Raw material in the manufacture of photographic chemicals; and
- Raw material in the production of azo dyes, dyes that contain nitrogen atoms at the core of their molecules

Acetaminophen is generally regarded as so safe that anyone can purchase it without a prescription. The drug is not entirely without its hazards, however. People who take large quantities of acetaminophen are at some risk for liver damage. That risk is especially great for anyone who drinks excessive amounts of alcohol. Alcohol by itself damages the liver and, in conjunction with acetaminophen, that damage can be serious. Some manufacturers now include a message

Words to Know

ANALGESIC A substance that relieves pain.

from simple beginning chemicals, or reactants.

SYNTHESIS A chemical reaction in which some desired chemical product is made

on the labels of their acetaminophen bottles about the risk of excessive ingestion of alcohol and acetaminophen.

FOR FURTHER INFORMATION

Kaplan, Tamara. "The Tylenol Crisis: How Effective Public Relations Saved Johnson & Johnson."

<http://www.personal.psu.edu/users/w/x/wxk116/tylenol/crisis.html> (accessed on September 14, 2005).

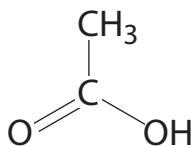
Keenan, Faith. "Blocking Liver Damage." *Business Week* (October 21, 2002): 147.

MedlinePlus Drug Information, Acetaminophen.

<http://www.nlm.nih.gov/medlineplus/druginfo/medmaster/a681004.html> (accessed on September 15, 2005).

Paracetamol Information Center.

<http://www.pharmweb.net/pwmirror/pwy/paracetamol/pharmwebpic.html> (accessed on September 15, 2005).



OTHER NAMES:	Ethanoic acid; methanecarboxylic acid; vinegar acid
FORMULA:	CH_3COOH
ELEMENTS:	Carbon, hydrogen, oxygen
COMPOUND TYPE:	Carboxylic acid (organic)
STATE:	Liquid
MOLECULAR WEIGHT:	60.05 g/mol
MELTING POINT:	16.6°C (61.9°F)
BOILING POINT:	117.9°C (244.2°F)
SOLUBILITY:	Soluble in water, alcohol, ether, acetone, benzene, and other organic solvents

KEY FACTS

Acetic Acid

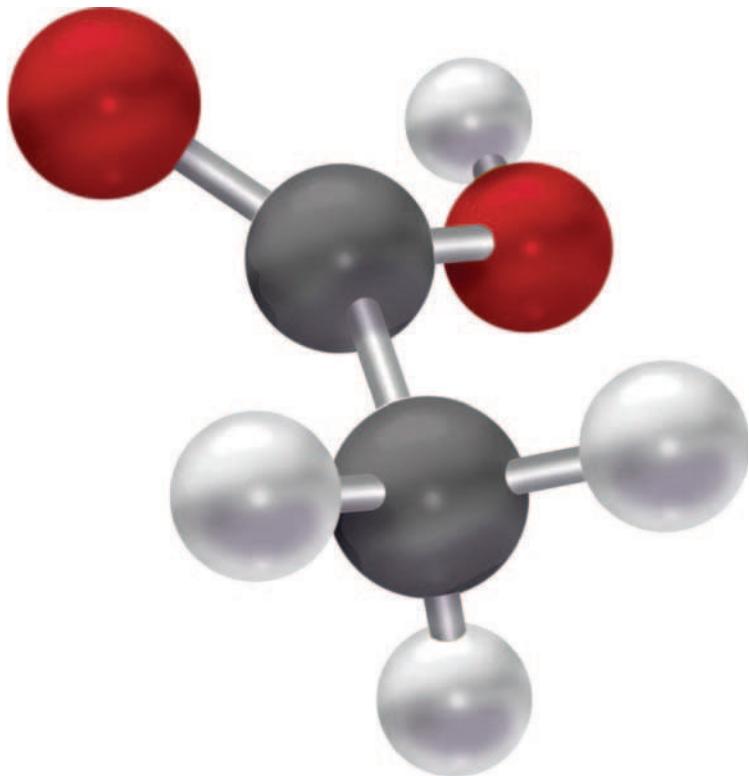
OVERVIEW

Acetic acid (uh-SEE-tik AS-id) is a clear, colorless liquid with a sharp odor. In its pure form, the compound is also known as glacial acetic acid. Acetic acid is the primary active ingredient of vinegar, which typically consists of about five parts of acetic acid mixed with 95 parts of water. The compound's name comes from the Latin word for vinegar, *acetum*.

Acetic acid, in the form of vinegar, has been known to humans for centuries. When fruit juices are allowed to stand for too long, or when they are fermented to make wine, vinegar forms. The use of vinegar as a condiment is mentioned a number of times in the Bible, and was described by the Greek natural philosopher Theophrastus (c. 372-c. 287 BCE). The first person to extract acetic acid from vinegar was the Muslim alchemist Jabir ibn Hayyan Geber (c. 721-815). The pure compound was not produced, however, for another ten centuries when the German chemist Georg Ernst Stahl (1660-1734) extracted acetic acid from vinegar in 1700 by distillation.

Acetic acid. Red atoms are oxygen; black atoms are carbon; and white atoms are hydrogen. The oxygen atom at top left shares a double bond with the nearby carbon atom.

PUBLISHERS RESOURCE GROUP



More than 1.4 million metric tons (1.5 million short tons) of acetic acid are produced in the United States annually. The largest fraction of that amount is used as a raw material in the manufacture of plastics.

HOW IT IS MADE

The most common method of making acetic acid is one developed by the Monsanto chemical corporation. In this process, synthesis gas (a mixture of carbon monoxide [CO] and hydrogen [H_2]) is heated over a catalyst of copper metal mixed with zinc oxide to make methanol (methyl alcohol; CH^3OH). The methanol is then treated with carbon monoxide (CO) to make acetic acid. Acetic acid can also be made by the fermentation of any material that contains sugar or some other carbohydrate. Although this method is of interest from a historical standpoint, it is not sufficiently efficient to use industrially.

Interesting Facts

Ancient Romans boiled fermented wine (vinegar) in lead pots to make a sweet syrup called sapa. The acetic acid in the vinegar dissolved a small amount of lead from which the pots were made, producing lead acetate

($\text{Pb}[\text{C}_2\text{H}_3\text{O}_2]$). When they used sapa as a sweetening agent in their foods, the Romans ingested the lead acetate which, over long periods of time, caused the lead poisoning from which so many of them died.

Researchers are constantly looking for new, more efficient, more environmentally-friendly methods for making acetic acid. In 2003, for example, chemists at the University of Southern California reported on a new method for making acetic acid directly from methane gas (CH_4) using a catalyst of palladium and sulfuric acid. Other researchers are looking for ways to oxidize the waste gases produced from industrial processes to acetic acid.

COMMON USES AND POTENTIAL HAZARDS

Acetic acid is probably best known to most people as vinegar. In this form, it is used as a condiment and a food preservative. The greatest volume of acetic acid is used, however, in a variety of chemical processes, especially the manufacture of plastic materials such as polyvinyl acetate (PVA), polyethylene terephthalate (PET), and cellulose acetate. A more recent use of acetic acid is in the manufacture of calcium magnesium acetate (CMA), a deicer. Traditionally, roads, highways, and airport runways have been treated with calcium chloride (CaCl_2) or some other salt to remove snow and ice. These compounds have serious environmental effects, however, and researchers have long been looking for alternatives that are as effective in removing snow and ice, but less harmful to the environment. CMA has been the most promising of these alternatives, and its production has produced a growing demand for acetic acid. Some other applications of acetic acid include:

- As a cleaning agent;
- In the manufacture of photographic materials;

Words to Know

ALCHEMY An ancient field of study from which the modern science of chemistry evolved.

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

DISTILLATION A process of separating two or more substances by boiling the mixture of which they are composed and condensing the vapors produced at different temperatures.

- For the production of a variety of organic compounds, such as those used in the manufacture of packaging materials, paints, adhesives and artificial fibers;
- As a fumigant (pesticide) to preserve fruits and grains;
- In the printing of textiles; and
- As an acidifier to improve the flow of oil from wells.

Dilute acetic acid in the form of vinegar is harmless and has been consumed by humans for centuries. Prolonged contact with the skin or eyes may, however, produce irritation of tissues and should be avoided. Concentrated forms of acetic acid pose more serious health risks, such as irritation of the gastrointestinal system, respiratory system, and eyes. Most people do not come into contact with the concentrated acid, and safety precautions are of importance only to individuals who handle the material in their work.

FOR FURTHER INFORMATION

“Acetic Acid.” Chemical of the Week.

<http://scifun.chem.wisc.edu/chemweek/AceticAcid/AceticAcid.html> (accessed on September 16, 2005).

“Acetic Acid and CMA.” <https://netfiles.uiuc.edu/mcheryan/www/cma.htm> (accessed on September 16, 2005).

“Acetic Acid and Derivatives.” *Kirk Othmer Encyclopedia of Chemical Technology*, 4th ed. Vol. 1. New York: John Wiley & Sons, 1991, pp. 121–175.

See Also Amyl acetate; butyl acetate; ethyl acetate; isoamyl acetate



OTHER NAMES:	Ethyne; ethine
FORMULA:	$\text{CH}=\text{CH}$
ELEMENTS:	Carbon, hydrogen
COMPOUND TYPE:	Alkyne (unsaturated hydrocarbon; organic)
STATE:	Gas
MOLECULAR WEIGHT:	26.04 g/mol
MELTING POINT:	-80.8°C (-113°F)
BOILING POINT:	-84°C (-120°F)
SOLUBILITY:	Slightly soluble in water and alcohol; soluble in acetone and benzene

KEY FACTS

Acetylene

OVERVIEW

Acetylene (uh-SET-ill-ene) is the simplest hydrocarbon, consisting of two carbon atoms joined to each other by a triple bond with their associated hydrogen atoms. It occurs as a colorless gas with a sweet odor when pure, but an unpleasant odor due to the presence of phosphine (PH_3) and/or arsine (AsH_3), with which it is often contaminated. Acetylene is a highly flammable gas that is also somewhat explosive. This property accounts for one of its most important uses, in torches that burn at very high temperatures.

Acetylene was discovered by the British chemist Edmund Davy (1785-1857) in 1836. Davy obtained the gas accidentally when he combined water with potassium carbide (KCH_2) while attempting to make potassium metal. He noted that the gas burned with a bright flame and thought it might be used as a source of illumination. That application was impractical, however, because of the high cost of potassium carbide. When the German chemist Frederich Wöhler (1800-1882) discovered the far less expensive calcium carbide



Acetyline. White atoms are hydrogen and black atoms are carbon. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

(CaC_2) in 1862, however, that problem was solved, and the demand for acetylene in lamps and other applications grew rapidly.

HOW IT IS MADE

The original method of making acetylene by the action of water on calcium carbide was abandoned when more efficient methods of synthesis became available. One of those methods, known as the Wulff process, involves the cracking of hydrocarbons from petroleum. In this process, liquid petroleum, which consists of many different kinds of hydrocarbons, is heated with very hot steam, causing the hydrocarbons to break apart ("crack") into smaller compounds. Acetylene is one of these compounds.

A later improvement on this procedure was developed by the German chemist Walter Reppe (1892–1969), sometimes called the Father of Acetylene Chemistry. Since Reppe worked at that time for the BASF chemical corporation, the process is also known as the BASF process of making acetylene. In this procedure, hydrocarbons from petroleum are treated with an oxidizing agent that causes them to break apart to form smaller compounds, one of which, again, is acetylene.

COMMON USES AND POTENTIAL HAZARDS

By far the most important use of acetylene is in the manufacture of other chemicals, such as vinyl chloride, vinylidene chloride, vinyl acetate, acrylonitrile, trichloroethylene

Interesting Facts

Acetylene was used as a source of illumination beginning in the early 1900s. In one type of acetylene lamp, water was allowed to drop on a solid chunk of calcium carbide at a controlled rate. The acetylene produced then passed into an ignition chamber, where it burned with a brilliant white light. Acetylene lamps were

used in a variety of settings, such as the source of light for bicycles and early motor cars and as a source of illumination for miners. Some small towns even used acetylene lamps as their major source of town lighting. Acetylene lamps continue to be a popular collectors item among antique dealers today.

and perchloroethylene, and the family of plastics known as the acrylates. (The name vinyl refers to the remnant of an acetylene molecule after one hydrogen atom has been removed: $\text{-CH}_2=\text{CH-}$) About 20 percent of all the acetylene produced is used for torches that produce very hot flames. In one such torch, the oxyacetylene torch, acetylene gas and oxygen are mixed and ignited at the tip of the torch. The combination of gases burns at a temperature of $3,000^{\circ}\text{C}$ to $3,500^{\circ}\text{C}$ ($5,500^{\circ}\text{F}$ to $6,300^{\circ}\text{F}$). Oxyacetylene torches are used to cut through metal and to weld two metals to each other. They can be used in very cold climates and even under water.

Acetylene is both very flammable and explosive. Any one who works with the compound or uses it in any form should know how to use the device that contains the gas. Acetylene also has the somewhat unusual chemical property of reacting with certain metals, such as copper and silver, to form highly explosive compounds known as acetylides. Lamps, torches, and other devices built to hold and dispense acetylene can not contain any of these metals. High concentrations of the gas also pose a health hazard to humans. It is classified as an asphyxiant, a gas that can produce disorientation, unconsciousness, and death when inhaled to excess.

Words to Know

OXIDATION A chemical reaction in which oxygen reacts with some other substance or, alternatively, in which some substances loses electrons to another substance, the oxidizing agent.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

FOR FURTHER INFORMATION

"Acetylene Properties, Purity and Packaging."

http://www.cfc.com/specgas_products/acetylene.htm

(accessed on September 16, 2005).

"The Environmental Impact of Acetylene Compared to Impact of Propane Chemtane 2." Chemtane.

http://www.chemtane2.com/environmental/cva_env_impact.html (accessed on September 16, 2005).

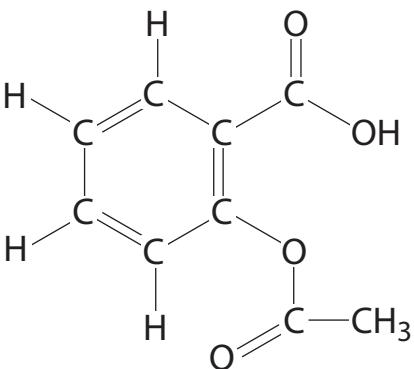
O'Hara, Kelly. "Chemistry Hall of Fame 1997 Acetylene." York University.

http://www.chem.yorku.ca/hall_of_fame/essays97/acetylene/acetylen.htm (accessed on September 16, 2005).

"Oxy acetylene Welding and Cutting." ESAB.

http://www.esabna.com/EUWeb/OXY_handbook/589oxy2_1.htm (accessed on September 16, 2005).

See Also Polyvinyl chloride



OTHER NAMES:

Aspirin; see Overview
for more names

FORMULA:

CH₃COOC₆H₄COOH

ELEMENTS:

Carbon, hydrogen,
oxygen

COMPOUND TYPE:

Carboxylic acid
(organic)

STATE:

Solid

MOLECULAR WEIGHT:

180.17 g/mol

MELTING POINT:

135°C (275°F;
decomposes)

BOILING POINT:

Not applicable

SOLUBILITY:

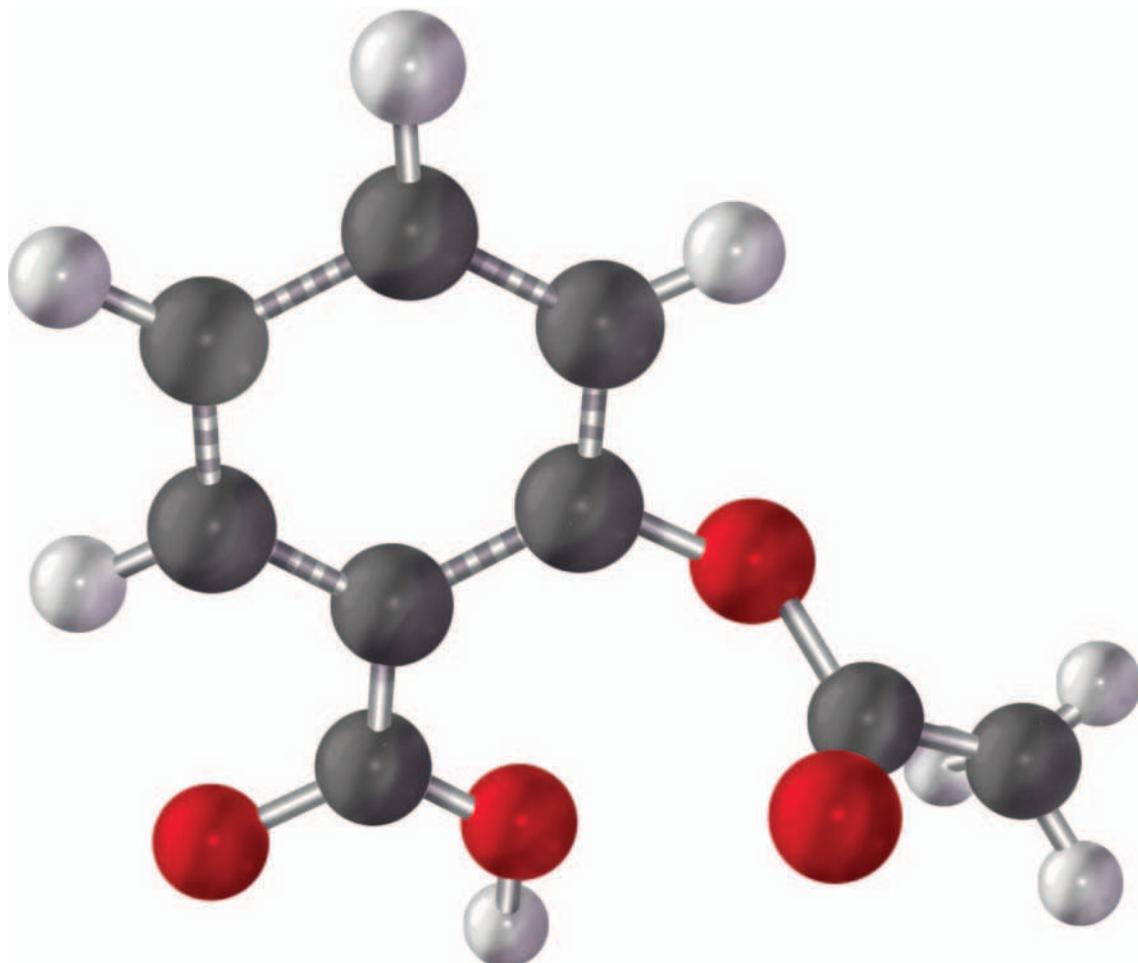
Soluble in water,
alcohol, ether,
chloroform

K E Y F A C T S

OVERVIEW

Acetylsalicylic acid (uh-SEE-till-sal-in-SILL-ik As-id, or uh-se-TEEL-sal-ih-SEEL-ik AS-id), more commonly known as aspirin, is the world's most commonly used therapeutic drug. By one estimate, about 137 million aspirin tablets are taken every day throughout the world. The drug is also known by other names including: *o*-acetoxybenzoic acid; 2-(acetyloxy)-benzoic acid; 2-carboxyphenyl acetate; and benzoic acid, 2-hydroxyacetate, in addition to about ten other systematic names and many common names.

The analgesic properties of willow tree bark, from which salicylic acid comes, have been known for well over 3,500 years. They were first described in Egyptian scrolls dating to about 1550 BCE and were later recommended by a number of ancient authorities, including the famous Greek physician Hippocrates (c. 460-370 BCE), the Roman encyclopedist Aulus Cornelius Celsus (c. 10 BCE-date of death unknown), the Roman philosopher Pliny the Elder (23 CE-CE), and the Greek physician Pedanius Dioscorides (40-90 CE).



Acetylsalicylic acid. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. Striped sticks indicate a benzene ring. PUBLISHERS RESOURCE GROUP

In the period from 1828 to 1829, the active ingredient in willow bark was first isolated by three individuals, the German pharmacist Johann Büchner (dates not available), the French chemist Henri Leroux (dates not available), and the Italian chemist, Raffaele Piria (1815–1865). Büchner gave the name salicin to the bitter-tasting yellow crystals extracted from willow bark after the Latin name for the willow tree, *Salix*. In 1853, the French chemist Charles Frederick Gerhardt (1816–1857) developed a method for reacting salicylic acid (the active ingredient in salicin) with acetic acid to make the first primitive form of aspirin.

Interesting Facts

- The name aspirin comes from a very old name for salicylic acid, spiraeic acid.
- After Hofmann's discovery, Bayer received a registered trademark for the name aspirin, meaning that no other company could use the name for

acetylsalicylic acid. When Germany was defeated in World War I (1914–1918), Allied forces seized the Bayer company's property and possessions and made the name available to all drug manufacturers.

For many years the way aspirin works in the body was not understood. Scientists now know that the compound's helpful effects come from its action on prostaglandins. Prostaglandins are hormone-like substances released by cells that are injured. They cause the body to release other substances that sensitize nerve endings to pain and start the healing process. Aspirin blocks prostaglandin production, thus relieving the sensation of pain and the inflammation that are the body's response to injury. Aspirin reduces fever by acting on the region of the brain that regulates body temperature and heart rate. Prostaglandins block the body's natural system for producing heat so that by blocking the release of prostaglandins, aspirin allows the regulation of body temperature to continue as usual. Aspirin's protection against heart attack and stroke occur because of its effect on one special type of prostaglandin, known as thromboxane A₂. Thromboxane A₂ promotes the accumulation of cells that takes place when a blood clot forms. By blocking or slowing down the production of thromboxane A₂, aspirin prevents the formation of blood clots and, hence, the probability of heart attack and stroke.

HOW IT IS MADE

The modern method for making aspirin was developed in 1897 by the German chemist Felix Hoffman (1868–1946), an employee of the German chemical manufacturer Bayer AG Chemical Works. In this procedure, phenol (C_6H_5OH) is treated with sodium hydroxide and carbon dioxide to make salicylic

Words to Know

ANALGESIC A substance that relieves pain. **THERAPEUTIC** Having healing properties.

acid. The salicylic acid is then reacted with acetic acid (CH_3COOH) to make acetylsalicylic acid, or aspirin. The preparation of aspirin by this procedure is quite simple and is often assigned to students in beginning high school and college chemistry classes. Aspirin tablets themselves include only acetylsalicylic acid, to which is added a small amount of water, starch and lubricant that act as a binder to hold the tablet together.

COMMON USES AND POTENTIAL HAZARDS

The exclusive use of aspirin is as a medicine. It has three important properties as a drug. It relieves pain, reduces inflammation, and reduces fever. In addition to its effectiveness in treating these medical symptoms, it is inexpensive and available in a variety of forms, including chewable tablets, extended-release formulations, effervescent tablets, and even in chewing gums. Aspirin is often prescribed in low, daily doses as a preventative measure for individuals at risk for heart attack and stroke.

While aspirin has many medical benefits, it is not without risk for some individuals. Some people are allergic to the compound and can not tolerate even a low dose. Such individuals experience a number of symptoms if they ingest high doses of aspirin, symptoms that include ringing in the ears, nausea, vomiting, dizziness, confusion, hallucinations, coma, seizures, rapid breathing, fever, and, in the most severe cases, death. Aspirin use is not recommended in children under the age of twelve who show symptoms of viral infections because it can lead to an extremely rare but deadly complication known as Reye's syndrome.

FOR FURTHER INFORMATION

Arnst, Catherine. "A Preemptive Strike against Cancer." *Business Week* (June 7, 2004): 48.

“Aspirin.” Plant Derived Drugs.

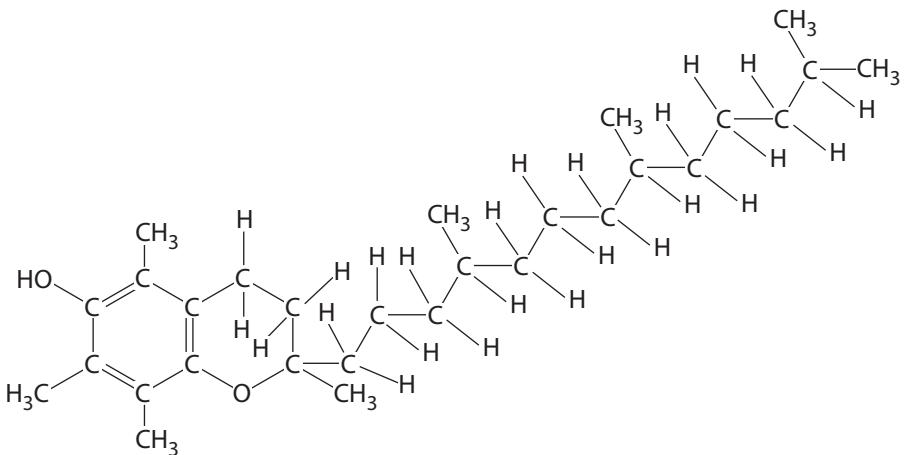
<http://phytomedical.com/Plant/Aspirin.asp> (accessed on September 17, 2005).

Jeffreys, Diarmuid. *Aspirin: The Remarkable Story of a Wonder Drug*. New York: Bloomsbury, 2004.

“Molecule of the Month: Aspirin.”

<http://www.bris.ac.uk/Depts/Chemistry/MOTM/aspirin/aspirin.htm> (accessed on September 17, 2005).

See Also Acetaminophen



OTHER NAMES:

See Overview.

FORMULA:

C₂₉H₅₀O

ELEMENTS:

Carbon, hydrogen,
oxygen

COMPOUND TYPE:

Organic

STATE:

Liquid

MOLECULAR WEIGHT:

430.69 g/mol

MELTING POINT:

2.5°C to 3.5°C (36°F to
38°F)

BOILING POINT:

350°C (660°F;
decomposes)

SOLUBILITY:

Insoluble in water;
soluble in alcohol,
ether, acetone, oils,
and most organic
solvents

K E Y F A C T S

OVERVIEW

Alpha-tocopherol (AL-fa toe-KOF-er-ol) is also known as 2,5,7,8-tetramethyl-2-(4,8,12-trimethyltridecyl)chroman-6-ol and 3,4-Dihydro-2,5,7,8-tetramethyl-2-(4,8,12-trimethyltridecyl)-2H-1-benzopyran-6-ol, as well as by many other systematic names. It is one of a family of compounds, the tocopherols, that makes up vitamin E. Other members of the family include the beta (β), gamma (γ), delta (δ), and epsilon (ϵ) tocopherols. All tocopherols share the same basic molecular structure, differing only in slightly different arrangements of methyl (CH₃) and hydroxyl (OH) groups. In terms of biological activity, α -tocopherol is the most important member of the tocopherol family. The tocopherols are all pale yellow, viscous oils found in a variety of plants, including almonds, mustard greens, green and red peppers, spinach, and sunflower seeds. The most important source for the vitamin is wheat germ.

Vitamin E was discovered in 1922 by two scientists at the University of California at Berkeley, Herbert McLean Evans



Alpha tocopheryl. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. PUBLISHERS RESOURCE GROUP

(1882-1971) and Katherine Scott Bishop (1889-1976). Evans and Bishop found that rats whose diet was deficient in α -tocopherol did not reproduce normally. Scientists are now aware that the primary function of vitamin D in the human body is as an antioxidant. An antioxidant is a substance that reacts with and prevents the harmful effects of chemical structures known as free radicals. A free radical is an atom or group of atoms with a single unpaired electron, making the atom or atoms unusually reactive. Free radicals are thought to be a primary cause of cell damage leading to cardiovascular disease, cancer, and aging.

Vitamin E was first produced synthetically (artificially) in 1938 by Swiss chemist Paul Karrer (1889-1971), who had been awarded the Nobel Prize in chemistry a year earlier for his studies of vitamin A, vitamin B_2 , and other important biological compounds.

HOW IT IS MADE

A variety of methods is now available for the synthesis (artificial production) of the tocopherols. In the most commonly used procedure, 2,3,5-trimethylhydroquinone is reacted with isophytol over one of many possible catalysts. A small amount of the vitamin is still obtained from natural sources, usually as the by-product in the treatment of one of its natural sources.

Interesting Facts

As a general rule, vitamins and minerals are equally potent (effective) whether they come from natural sources or synthetic procedures. Vitamin E is an exception to that general rule. Synthetic forms of the vitamin are generally less potent than natural forms. The reason for this fact is that natural vitamin E exists in only one stereo isomeric form. Stereo isomers are two or more forms of a compound with the same kind and number of atoms. But the atoms in stereoisomers have some

what different orientations in space. These slight differences in geometric shape are responsible for slight differences in biological activity. Natural vitamin E consists of only one stereo isomer, which is very active biologically. Synthetic vitamin E consists of two stereoisomers, only one of which (the one that occurs in natural vitamin E) is biologically active. So synthetic vitamin E tends to be about half as active biologically as natural vitamin E.

COMMON USES AND POTENTIAL HAZARDS

Vitamin E deficiency disorders are rare in humans. The vast majority of people get all the vitamin E they need in their daily diets. Three groups of people may, however, require vitamin E supplements. The first group consists of individuals who are unable to absorb fat in their diets, so that the vitamin E they ingest is immediately eliminated from their bodies. The second group consists of a small fraction of people who have genetic disorders that make it impossible for them to utilize α -tocopherol. The third group of individuals consist of premature and very young babies. All babies are born deficient in vitamin E, but they normally overcome that deficiency in the first few weeks of their lives. Premature babies and some babies born at term may experience a variety of health problems as a result of vitamin E deficiency. Lack of normal reflexes, inability to orient itself, muscle weakness, and loss of balance are typical symptoms of vitamin E deficiency in babies. Deficiency problems for individuals in all three groups can be overcome by taking supplementary vitamin E.

Words to Know

CARCINOGEN A chemical that causes cancer in humans or other animals.

VISCOUS Describes a syrupy liquid that flows slowly.

In addition to its use as a vitamin supplement for normal individuals and those at risk for vitamin E deficiency, the tocopherols have a few other uses:

- In the curing of meat to block the action of nitrosamines, a group of compound that occurs naturally in meats and may be carcinogenic;
- As an additive to animal feed to replace vitamins lost during feed processing; and
- As a food additive in vegetable oils and shortening to prevent oxidation (spoilage).

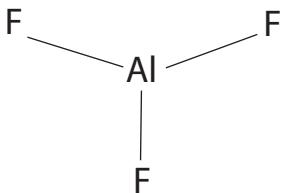
FOR FURTHER INFORMATION

Challem, Jack, and Melissa Diane Smith. *Basic Health Publications User's Guide to Vitamin E: Don't Be a Dummy: Become an Expert on What Vitamin E Can Do for Your Health*. North Bergen, NJ: Basic Health Publications, 2002.

Packer, Lester, and Carol Colman. *The Antioxidant Miracle: Put Lipoic Acid, Pycogenol, and Vitamins E and C to Work for You*. New York: Wiley, 1999.

"Vitamin E." National Institutes of Health, Office of Dietary Supplements.
<http://ods.od.nih.gov/factsheets/vitamine.asp> (accessed on September 17, 2005).

"Vitamin E (tocopherol)." Vitamin and Health Supplements Guide.
<http://www.vitaminsupplements.org/vitaminE.php> (accessed on September 17, 2005).



OTHER NAMES:	Aluminum trifluoride
FORMULA:	AlF_3
ELEMENTS:	Aluminum, fluorine
COMPOUND TYPE:	Inorganic salt
STATE:	Solid
MOLECULAR WEIGHT:	83.98 g/mol
MELTING POINT:	1291°C (2356°F; begins to sublime above 1250°C [2280°F])
BOILING POINT:	Not applicable
SOLUBILITY:	Slightly soluble in cold water, soluble in hot water; insoluble in alcohol, acetone, and most organic solvents

KEY FACTS

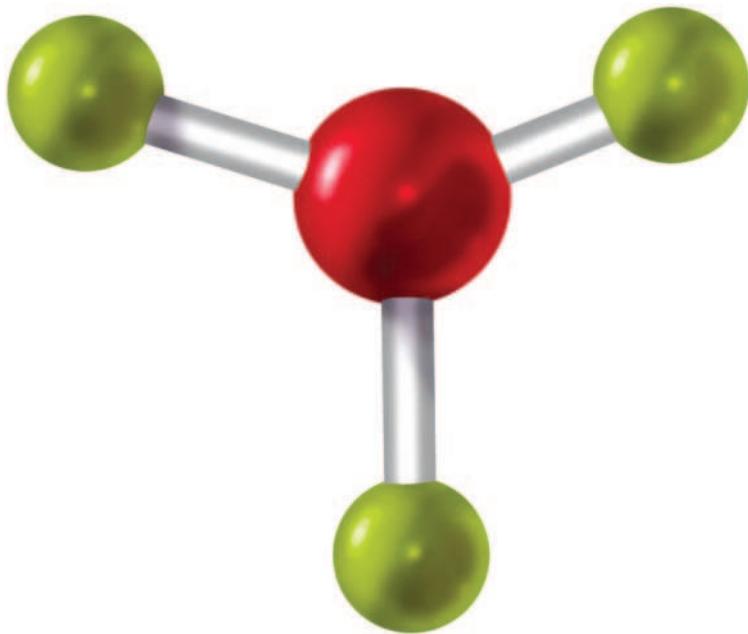
Aluminum Fluoride

OVERVIEW

Aluminum fluoride (uh-LOO-min-um FLOR-ide) is a highly stable compound that occurs as a white crystalline solid. It resists the action of even strong solvents, such as hot concentrated sulfuric acid. The compound often occurs as a hydrate containing one or more molecules of water of hydration. The most common of these hydrates has the chemical formula $\text{AlF}_3 \cdot 3.5\text{H}_2\text{O}$, meaning that for every two molecules of aluminum fluoride in a crystal, there are seven molecules of water. The major uses of aluminum fluoride are in a variety of applications in the chemical industry.

HOW IT IS MADE

Three methods are available for the preparation of aluminum fluoride commercially. In the first method, alumina trihydrate (aluminum hydroxide; $\text{Al}(\text{OH})_3$) is treated with hydrofluoric acid (HF). A second method of preparation is almost identical except that it takes place in the dry state



Aluminum fluoride. Red atom is aluminum and green atoms are fluorine. PUBLISHERS RESOURCE GROUP

between dry alumina trihydrate and gaseous hydrogen fluoride. The third method involves the addition of fluorosilicic acid (H_2SiF_6) to alumina trihydrate. The last of these methods is commercially attractive because fluorosilicic acid is obtained inexpensively as a by-product in the production of fertilizers and phosphoric acid.

COMMON USES AND POTENTIAL HAZARDS

The primary use of aluminum fluoride is in the production of aluminum metal. In that process, aluminum is extracted from one of its compounds (usually aluminum oxide) by passing an electric current through the molten (melted) compound. The addition of aluminum fluoride to the raw materials used in the process reduces the temperature at which the reaction occurs and improves the conductivity of the molten compound.

Some other uses of aluminum fluoride include:

- As a catalyst in the synthesis of organic compounds;
- To suppress the process of fermentation in the wine- and beer-making industries;

Interesting Facts

Controversy exists about the health hazards of aluminum fluoride. On the one hand, the compound would appear to pose a relatively minor risk when ingested because it is so insoluble in water and other solvents. If it does not dissolve in the body, it can not get into the blood

stream. On the other hand, it can cause health problems if inhaled or deposited on the skin. Opponents of water fluoridation argue that aluminum fluoride is sufficiently toxic to humans that it should not be added to public water supplies.

- In the manufacture of ceramics and glass; and
- As a flux in the preparation of glazes, enamels and in metallurgical processes.

Most authorities agree that aluminum fluoride does not pose a health hazard when ingested in moderate amounts because of its low solubility. The compound may pose a threat if inhaled or deposited on the skin, however. In modest amounts, it may produce skin and eye irritations that range from moderate to severe; irritation of the nose, throat, and lungs, that may include nosebleeds; and asthma-like

Words to Know

FLUX A material that lowers the melting point of another substance or mixture of substances or that is used in cleaning a metal.

HYDRATE A chemical compound formed when one or more molecules of water is added physically to the molecule of some other substance.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

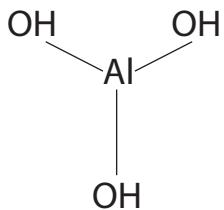
WATER OF HYDRATION Water that has combined with a compound by some physical means.

ALUMINUM FLUORIDE

symptoms with coughing, shortness of breath, and tightness in the chest. Long-term exposure may result in weakness of the bones and stiffening of the joints.

FOR FURTHER INFORMATION

- “Aluminum Fluoride.” ChemExper Chemical Directory.
<http://www.chemexper.com/chemicals/supplier/cas/7784-18-1.html> (accessed on September 18, 2005).
- “Aluminum Fluoride.” Chemical Land 21.
<http://www.chemicalland21.com/industrialchem/inorganic/ALUMINUM%20FLUORIDE.htm> (accessed on September 18, 2005).
- “Aluminum Fluoride.” Solvay Chemicals.
http://www.solvaychemicals.us/pdf/Inorganic_Fluorides/ALF.pdf (accessed on September 18, 2005).



OTHER NAMES:	See Overview.
FORMULA:	$\text{Al}(\text{OH})_3$
ELEMENTS:	Aluminum, oxygen
COMPOUND TYPE:	Inorganic base
STATE:	Solid
MOLECULAR WEIGHT:	78.00 g/mol
MELTING POINT:	Not applicable; loses water on heating to become aluminum oxide (Al_2O_3)
BOILING POINT:	Not applicable
SOLUBILITY:	Insoluble in water; soluble in strong acids and bases

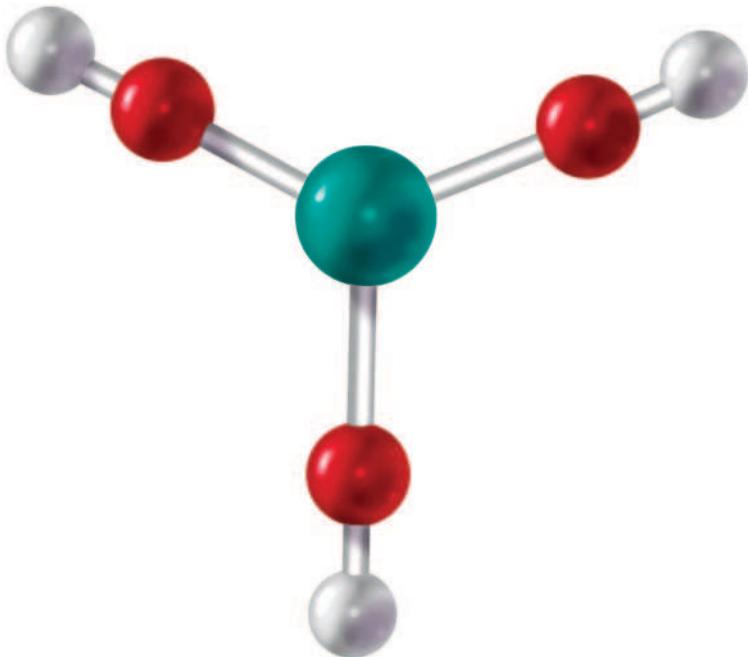
KEY FACTS

OVERVIEW

Aluminum hydroxide (uh-LOO-min-um hi-DROK-side) is also known as alumina hydrate, aluminum hydrate, aluminum trihydrate, hydrated alumina, hydrated aluminum oxide, and aluminum white. It is a white crystalline powder that occurs in nature as the mineral gibbsite, one of the three minerals (along with boehmite, diaspore) that makes up bauxite. Bauxite is a rocky material from which the metal aluminum is extracted commercially. Aluminum hydroxide is perhaps best known to the average person as an ingredient in antacids, compounds for the treatment of upset stomach, acid indigestion, and related symptoms. It is found in products such as Maalox®, Mylanta®, and Gelusil®. A closely related chemical compound, aluminum hydroxychloride ($\text{Al}_2\text{Cl}(\text{OH})_5$) is the active ingredient in most deodorants and anti-perspirants.

HOW IT IS MADE

Most aluminum hydroxide is produced by treating bauxite with a hot, concentrated solution of sodium hydroxide



Aluminum hydroxide. Red atoms are oxygen; white atoms are hydrogen; and turquoise atom is aluminum. PUBLISHERS RESOURCE GROUP

(NaOH). As the solution cools, aluminum hydroxide settles out of the reaction mixture as a gelatinous (jelly-like) precipitate, which can then be dried and saved as a white powder. The aluminum hydroxide can also be precipitated by passing carbon dioxide gas through the solution.

COMMON USES AND POTENTIAL HAZARDS

Aluminum hydroxide is an effective antacid because it is, chemically, a base. As such, it reacts with stomach acid (hydrochloric acid; HCl), to reduce the symptoms of heartburn, upset stomach, acid indigestion, and gastritis, an inflammation of the stomach lining. The compound can also be used to treat peptic ulcers on a long-term basis. A peptic ulcer is an open sore in the lining of the stomach or the first part of the small intestine.

Aluminum hydroxide also has a number of industrial applications, such as:

- A raw material for the preparation of other aluminum compounds;
- An additive for cosmetics, paper, plastics, and rubber to give the final product more bulk;

Interesting Facts

Aluminum hydroxide is one of the most common adjuvants used in vaccines today. An adjuvant is a substance added to a vaccine to increase the efficiency with which it

prevents disease. The mechanisms by which adjuvants such as aluminum hydroxide achieve this objective are not always known.

- In abrasives designed for use with plastics and brass;
- An additive for glass products that adds strength and resistance to attack by chemicals and weathering and to heat shock;
- As the world's most popular flame retardant.

Studies have not shown that aluminum hydroxide poses health risks of any kind to human beings.

FOR FURTHER INFORMATION

"Aluminum Hydroxide." Chemical Land 21.

<http://www.chemicaland21.com/industrialchem/inorganic/aloh3.htm> (accessed on September 18, 2005).

"Aluminum Hydroxide." Drugs.com.

http://www.drugs.com/MTM/aluminum_hydroxide.html (accessed on September 18, 2005).

Words to Know

ADJUVANT A substance added to a vaccine to increase the efficiency with which it prevents disease.

ANTACID A medicine used for the treatment of upset stomach, acid indigestion, and related symptoms.

PRECIPITATE A solid material that settles out of a solution, often as the result of a chemical reaction.

ALUMINUM HYDROXIDE

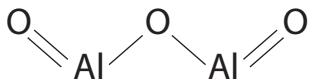
“Aluminum Hydroxide Resources.” drugs@mongabay.com.

<http://names.mongabay.com/drugs/ingredients/>

ALUMINUM_HYDROXIDE.html (accessed on September 18, 2005).

Malakoff, David. “Public Health: Aluminum Is Put on Trial as a Vaccine Booster.” *Science* (May 26, 2000): 1323.

See Also Aluminum oxide

**OTHER NAMES:**

Alumina

FORMULA: Al_2O_3 **ELEMENTS:**

Aluminum, oxygen

COMPOUND TYPE:

Metallic oxide

STATE:

Solid

MOLECULAR WEIGHT:

101.96 g/mol

MELTING POINT: 2072°C (3762°F)**BOILING POINT:** 2980°C (5396°F)**SOLUBILITY:**

Insoluble in water and most organic solvents; slowly soluble in basic solutions.

**K
E
Y
F
A
C
T
S**

Aluminum Oxide

OVERVIEW

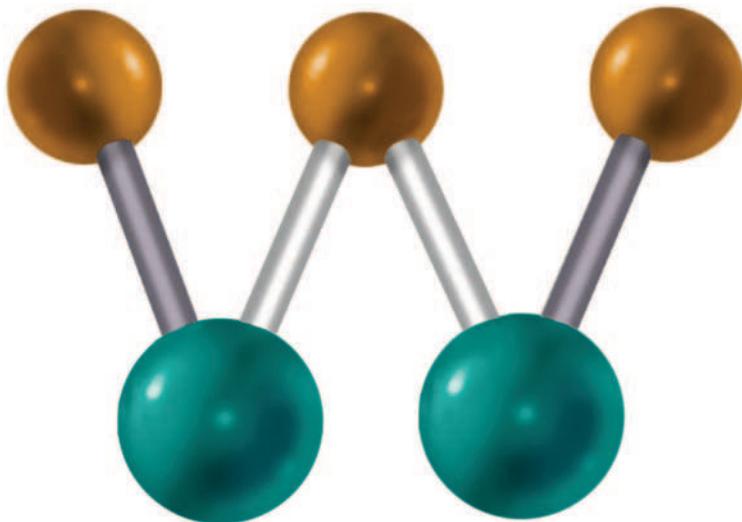
Aluminum oxide (uh-LOO-min-um OK-side) is white crystalline powder that occurs in nature in a variety of minerals, including boehmite, bayerite, corundum, diasporite, and gibbsite. Corundum is second hardest naturally occurring mineral. Only diamond is harder. Aluminum oxide occurs in a variety of chemical forms in a variety of gemstones, including chrysoberyl, ruby, sapphire, and spinel. The color of these gemstones is a result of impurities, such as chromium (in the case of ruby) and iron and titanium (in the case of sapphire). The colors may also vary depending on the kind and amount of each impurity.

Aluminum oxide's commercial uses depend not only on its hardness, but also on its high melting point and its low electrical conductivity. The compound is also non-combustible and resistant to attack by most solvents and other chemical agents.

HOW IT IS MADE

Aluminum oxide is produced by washing the rocky material bauxite with a hot solution of sodium hydroxide (NaOH).

Aluminum oxide. Turquoise atoms are aluminum and orange atoms are oxygen. Gray sticks indicate double bonds.
PUBLISHERS RESOURCE GROUP



The aluminum hydroxide ($\text{Al}(\text{OH}_3)$) that is produced in this reaction is then heated to drive off water, producing aluminum oxide. Wastes from coal mining operation are also treated to extract the aluminum sulfate ($\text{Al}_2(\text{SO}_4)_3$) they contain. The aluminum sulfate is converted to aluminum hydroxide which, again, is heated to produce aluminum oxide.

COMMON USES AND POTENTIAL HAZARDS

The primary use of aluminum oxide is the manufacture of aluminum metal. When an electric current is passed through molten (melted) aluminum oxide, the compound breaks down to form aluminum metal and oxygen gas. The method is called the Hall process after the American chemist Charles Martin Hall who invented it.

Aluminum oxide is also widely used as an abrasive. An abrasive is a very hard material used to grind, polish, sand, scour, scrub, smooth or polish some other material. Among the products that include aluminum oxide as an abrasive are emery boards, sandpaper, grinding and polishing wheels and belts, lens grinding devices, and gem polishing wheels.

The high melting point of aluminum oxide also makes it a good refractory product. A refractory material is one that does not melt easily, making it suitable for lining the inside

Interesting Facts

- The hardness of materials is measured on the Mohs scale, named after the German mineralogist Frederick Mohs (1773–1839), who suggested the system. The scale runs from 1 (the softest known natural material, talc) to 10 (the hardest known natural material, diamond). Corundum has a hardness of 8.8 on the Mohs scale.
- The invention of a method for making aluminum metal from aluminum oxide by the 23-year-old American chemist Charles Martin Hall (1863–1914) in 1886 resulted in a reduction in the price of aluminum metal from about twelve dollars a pound to less than a dollar a pound.

of furnaces or the manufacture of glass and ceramic materials that will not melt when exposed to very high temperatures. Some other uses of aluminum oxide include:

- In finely-divided form, as the packing material in chromatographic columns. Chromatography is a process by which individual components of a mixture are separated from each other by passing them through a tube filled with some absorbent material (such as aluminum oxide).
- As a catalyst in many industrial chemical reactions;
- In the paper-making process, as a filler that adds body to the final product;
- For the production of artificial gemstones;
- As a food additives, where it acts as a dispersant, a substance that keeps a product from clumping together in a package; and
- As the internal coating on frosted light bulbs.

The U.S. Occupational Safety and Health Administration (OSHA) classifies aluminum oxide as a *nuisance dust* in the workplace. A nuisance dust is one for which no serious harmful effects have been identified as long as its release is kept under control. It may, however, cause unpleasant

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure

CHROMATOGRAPHY A process by which a mixture of substances passes through a column consisting of some material that causes the individual components in the mixture to separate from each other.

DISPERSANT A substance that keeps another substance from clumping together or becoming lumpy.

REFRACTORY Does not melt easily; able to withstand high temperatures.

SOLVENT A substance that is able to dissolve one or more other substances.

symptoms, such as irritation of the skin, eyes, and lungs if present in unusually large amounts. The simplest treatment for problems related to nuisance dusts is to move away from the contaminated area where an adequate supply of fresh air is available.

FOR FURTHER INFORMATION

"Aluminum Oxide, Al₂O₃." Accuratus.

<http://www.accuratus.com/alumox.html> (accessed on September 19, 2005).

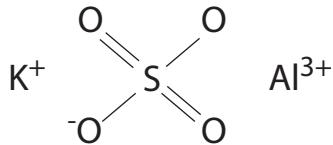
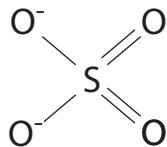
Frost, Randall. "Abrasives." *Gale Encyclopedia of Science*. Edited by K. Lee Lerner and Brenda Wilmoth Lerner. 3rd ed. Vol. 1. Detroit, Mich.: Gale, 2004.

"The Mineral Corundum." Amethyst Galleries.

<http://mineral.galleries.com/minerals/oxides/corundum/corundum.htm> (accessed on September 19, 2005).

Misra, Chanakya. *Industrial Alumina Chemicals*. Washington, D.C.: American Chemical Society, 1986.

See Also Aluminum hydroxide

**OTHER NAMES:**

See Overview.

FORMULA:

$\text{KAl}(\text{SO}_4)_2$

ELEMENTS:

Potassium, aluminum, sulfur, oxygen

COMPOUND TYPE:

Inorganic salt

STATE:

Solid

MOLECULAR WEIGHT:

258.20 g/mol

MELTING POINT:

92.5°C (198°F; loses water of hydration beginning at 64.5°C (148°F)

BOILING POINT:

200°C (392°F; loses all water of hydration)

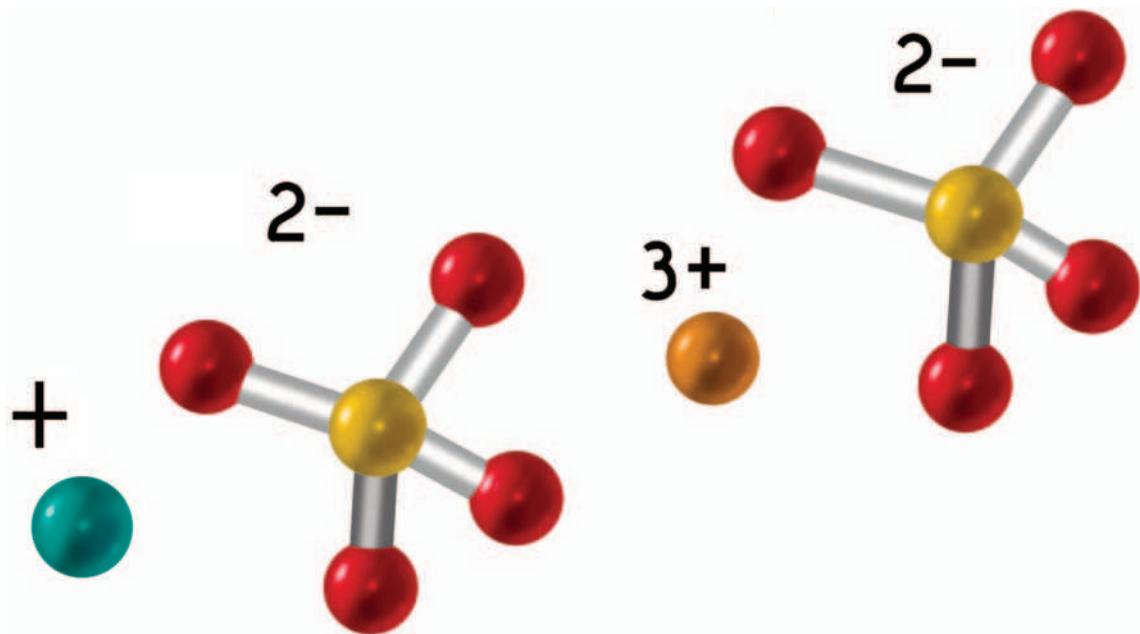
SOLUBILITY:

Soluble in cold water; very soluble in hot water; insoluble in alcohol, ether, acetone, and other organic solvents

K E Y F A C T S**OVERVIEW**

Aluminum potassium sulfate (uh-LOO-min-um po-TASS-see-um sul-fate) is also known as aluminum potassium sulfate dodecahydrate, potash alum, potassium alum, and kalinite. It normally occurs in the form of the dodecahydrate, meaning that each molecule of the compound is associated with twelve molecules of water. The formula of the hydrate is $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$. In this form, it occurs as white odorless crystals. The compound gradually loses its water of hydration when heated, giving up the first nine molecules of water at 64.5°C (148°F), and the remaining three molecules of water at about 200°C (392°F).

Aluminum potassium sulfate belongs to a family of compounds known collectively as the alums. The term *alum* refers to a double salt that consists of aluminum, the sulfate group (SO_4), and one other metal. The presence of two metals, aluminum plus one other metal, accounts for the name *double salt*. Other common alums are aluminum ammonium sulfate and aluminum sodium sulfate.



Aluminum potassium sulfate.
Red atoms are oxygen; yellow atoms are sulfur; turquoise atom is potassium; and orange atom is aluminum. Gray sticks indicate double bonds.

PUBLISHERS RESOURCE GROUP

The alums were known as far back as ancient Egypt and China, where they were used as deodorants. Alums are effective for this purpose because they are astringents, substances that cause tissues to shrink or contract, thus reducing the tendency of sweat glands to produce perspiration. As astringents, alums were also used in the field of medicine to treat wounds and prevent bleeding. Until the nineteenth century, however, chemists did not recognize that the substance they knew as alum was actually a variety of different compounds.

HOW IT IS MADE

Aluminum potassium sulfate occurs naturally in the form of the minerals alunite and kalinite. Where these minerals are available, the compound can be mined and purified to obtain potash alum. Where the minerals are not available, the compound can be produced synthetically by combining aqueous aluminum sulfate with aqueous potassium sulfate. The two compounds react with each other in solution to form the double salt, aluminum potassium sulfate, which can then be extracted by allowing the solutions to evaporate, during which the desired compound crystallizes out.

Interesting Facts

- During the fifteenth century, the Vatican achieved control of the alum industry in Europe. When King Henry VIII of England quarreled with Pope Clemens VII a century later over his desire to marry a

second time, the pope cut off supplies of alum to Great Britain. Since alum was an essential product in the dyeing of clothes, the English clothing industry rapidly fell into dire circumstances.

COMMON USES AND POTENTIAL HAZARDS

One of aluminum potassium sulfate's major uses is in the dyeing of fibers and fabrics, where it is employed as a mordant. A mordant is a substance that reacts with a dye, helping it attach more permanently to a fiber or fabric. Aluminum potassium sulfate has also been used in the paper-making industry for many centuries, where it has a variety of applications. For example, it can be used to give paper a tough, shiny surface or to increase the intensity of inks, paints, and dyes used on the paper. Some water treatment plants also use aluminum potassium sulfate in their purification systems. The compound is added to water, where it combines with colloidal particles suspended in water to form larger clumps, which then settle out of the water. Other uses of aluminum potassium sulfate include:

- As a food additive, used to control the acidity of the product;
- In the manufacture of matches;
- For the waterproofing of fabrics;
- In the tanning of leather;
- In the manufacture of deodorants;
- To add hardness and toughness to cement;
- In the production of fireworks;

Words to Know

AQUEOUS SOLUTION A solution that consists of some material dissolved in water.

ASTRINGENT A material triggers a loss of water from tissue, thereby causing the tissues to shrink and contract.

COLLOID A mixture consisting of tiny particles suspended, but not dissolved, in water.

DOUBLE SALT A salt that includes aluminum and one other metal, such as potassium.

MORDANT A substance that reacts with a dye so that it attaches to a fiber or fabric more easily.

- As an astringent in medical treatments; and
- In the preparation of other compounds of aluminum in the chemical industry.

FOR FURTHER INFORMATION

“Aluminum Potassium Sulfate Resources.”

http://names.mongabay.com/drugs/ingredients/ALUMINUM_POTASSIUM_SULFATE.html (accessed on September 18, 2005).

Hart Davis, Adam. “Thomas Chaloner and His Astonishing Alum Industry.” ExNet.

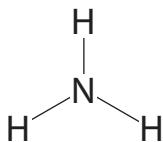
<http://www2.exnet.com/1995/12/18/science/science.html#LINKS BACK>
(accessed on September 18, 2005).

“Material Safety Data Sheet: Aluminum Potassium Sulfate Dodecahydrate.” Department of Chemistry, Iowa State University.

<http://avogadro.chem.iastate.edu/MSDS/AIKSO4.htm>
(accessed on September 18, 2005).

“Potassium Aluminum Sulfate.” *Gale Encyclopedia of Science*. Edited by K. Lee Lerner and Brenda Wilmoth Lerner. 3rd ed. Vol. 5. Detroit, Mich.: Gale, 2004.

See Also Potassium Sulfate

**OTHER NAMES:**

None

FORMULA: NH_3 **ELEMENTS:**

Nitrogen, hydrogen

COMPOUND TYPE:

Inorganic base

STATE:

Gas

MOLECULAR WEIGHT:

17.03 g/mol

MELTING POINT: -77.7°C (-108°F)**BOILING POINT:** -33.35°C (-28.03°F)**SOLUBILITY:**

Very soluble in cold water; soluble in alcohol, ether, and many organic solvents

**K
E
Y
F
A
C
T
S**

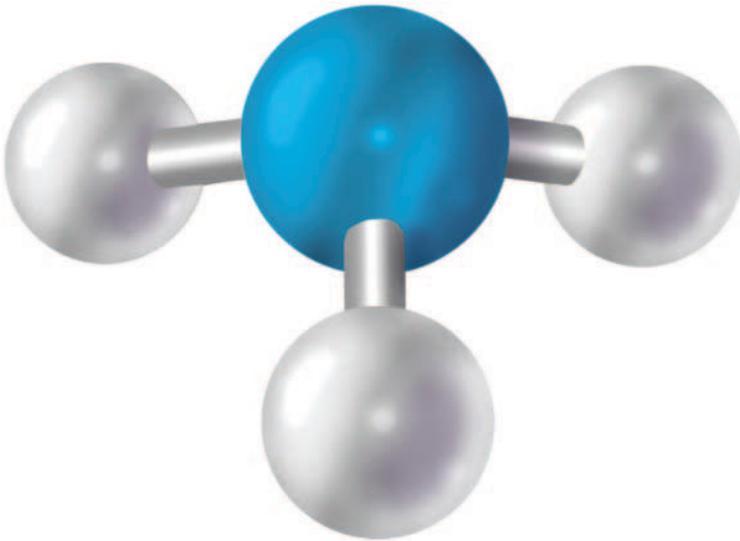
Ammonia

OVERVIEW

Ammonia (uh-MOH-nyah) is a colorless gas with a strong, suffocating odor. It was present in the primordial (original) atmosphere of the Earth. Scientists believe that it may have been the source of nitrogen for the earliest forms of life. Ammonia was the first chemical compound to be found in interstellar space, the space between stars. It is a major component of the atmosphere of many planets in our solar system.

Early chemists learned to produce ammonia from animal parts, such as the horns of deer. But it was the English chemist and physicist Joseph Priestley (1733-1804) who first collected and studied the pure gas. In 1785, the French chemist Claude-Louis Berthollet (1748-1822) determined the correct chemical formula for the gas, NH_3 .

In 2004, American companies produced 10,762,000 metric tons (11,863,000 short tons) of ammonia, making it the tenth highest-volume chemical made in the United States. Only ten years earlier, it had ranked number 5 on



Ammonia. Blue atom is nitrogen and white atoms are hydrogen. PUBLISHERS RESOURCE GROUP

the list of all chemicals produced by volume. About 90 percent of all the ammonia used in the United States goes to the production of fertilizers.

HOW IT IS MADE

Ammonia is produced naturally by the action of certain types of bacteria on nitrogen found in the atmosphere. It is also formed during the decay of dead organisms.

Until the end of the nineteenth century, ammonia was produced commercially by the cyanamide process. Calcium carbide (CaC_2), nitrogen gas (N_2), and steam were reacted with each other to produce ammonia.

In the early 1900s, the German chemist Fritz Haber (1868-1934) developed a method for making ammonia directly from its elements, nitrogen and hydrogen. The two gases are combined with each other at high temperature (400°C to 650°C; 750°F to 1200°F) and pressure (200 to 400 atmospheres; 3,000 to 6,000 pounds per square inch) over a catalyst made of finely-divided iron. Haber's process was later refined and improved by German chemical engineer Carl Bosch (1874-1940). Haber and Bosch both won Nobel Prizes in chemistry for their work on the production of ammonia. The Haber-Bosch process remains the most common form of ammonia production in many countries, including the United

Interesting Facts

- Ammonia's name comes from an ancient Egyptian practice conducted at the temple of the sun god Amon (or Ammon) near Karnak. Camel dung burned at the temple gave off a strong odor (ammonia) and left behind a white crystalline substance on the ground. The white substance was given the name of sal ammoniac, or salt of Amon, and the gas itself later became known as ammonia.
- The Haber-Bosch process was developed largely because of Germany's need for explosives in World War I. Ammonia gas is converted to nitric acid, which, in turn, is used in making sodium and potassium nitrate, major components of explosives. Fritz Haber believed that it was his patriotic duty to contribute to the German war effort in whatever way he could, which led to his development of a new method for making ammonia.

States. Small amounts of ammonia are also produced during the process by which soft coal is converted to coke.

Ammonia is a natural product of metabolism in all animals. When proteins break down, the nitrogen they contain is converted, in part, to ammonia. The ammonia is then converted to urea, which is excreted in the urine.

COMMON USES AND POTENTIAL HAZARDS

Ammonia is used in a variety of forms as a fertilizer. It can be liquified or dissolved in water and sprayed on land, or it can be converted into any one of a number of compounds, such as ammonium nitrate, ammonium phosphate, or ammonium sulfate. In these forms, it is spread as dry granules on the land. Urea, made from ammonia and carbon dioxide, is also used as a feed supplement for cattle. Plants and animals use the

nitrogen in ammonia and its compounds to synthesize new proteins that contribute to their growth and development.

The next largest use of ammonia is in the synthesis of nitric acid (HNO_3). In a process developed by the German chemist Wilhelm Ostwald (1853–1932), ammonia, oxygen, and water are reacted together in a series of steps that results in the formation of nitric acid. Nitric acid, the thirteenth most important chemical in the United States in terms of productions, has a number of important uses, including the manufacture of explosives. Like the Haber-Bosch process, the Ostwald process contributed to the success experienced by Germany during World War I.

In addition to its use in the manufacture of fertilizers and explosives, smaller amounts of ammonia are used:

- As a refrigerant;
- In the manufacture of plastics;
- As a raw material in the manufacture of other nitrogen-containing chemicals;
- In the production of dyes;
- As a rocket fuel;
- For the neutralization of acids during the refining of petroleum;
- In order to produce specialized types of steel; and
- As a nutrient in yeast cultures in food processing operations.

Both gaseous and liquid ammonia pose moderate health hazards to those who come into contact with them. For example, farmers who handle liquid ammonia risk the possibility of painful blistering of the skin or damage to the mucous membranes if they come into contact with the fertilizer. Ammonia fumes can irritate the mouth, nose, and throat, causing coughing and gagging responses. Higher levels of exposure may irritate the lungs, resulting in shortness of breath and producing headaches, nausea, and vomiting. Very high exposures can cause a buildup of fluid in the lungs that can result in death. Since ammonia is a common ingredient of many household products, everyone should be aware of its health risks, although the threat posed by such products is, in fact, very small.

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

METABOLISM The process including all of the chemical reactions that occur in cells by which fats, carbohydrates, and other compounds are broken down to produce energy and the compounds needed to build new cells and tissues.

MUCOUS MEMBRANES Tissues that line the moist inner lining of the digestive, respiratory, urinary and reproductive systems.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

FOR FURTHER INFORMATION

“Ammonia.” Masterliness.

<http://www.masterliness.com/a/Ammonia.htm>
(accessed on September 19, 2005).

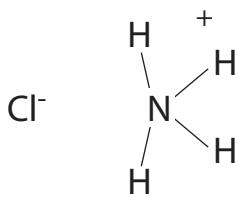
Buechel, K. H., et al. Industrial Inorganic Chemistry. New York: VCH, 2000, pp. 29-43.

“The Facts about Ammonia.” New York State Department of Health.
[http://www.health.state.ny.us/nysdoh/bt/chemical_terrorism/
docs/ammonia_general.pdf](http://www.health.state.ny.us/nysdoh/bt/chemical_terrorism/docs/ammonia_general.pdf)
(accessed on September 19, 2005).

“Toxicological Profile for Ammonia.” Agency for Toxic Substances and Disease Registry.
<http://www.atsdr.cdc.gov/toxprofiles/tp126.html>
(accessed on September 19, 2005).

“Uses and Production of Ammonia (Haber Process).” Ausetute.
<http://www.ausetute.com.au/haberpro.html>
(accessed on September 19, 2005).

See Also Ammonium Chloride; Ammonium Hydroxide; Ammonium Nitrate; Ammonium Sulfate; Nitric Acid; Potassium Nitrate; Urea



OTHER NAMES:

Ammonium muriate;
sal ammoniac;
salmiac

FORMULA:

NH_4Cl

ELEMENTS:

Nitrogen, hydrogen,
chlorine

COMPOUND TYPE:

Inorganic salt

STATE:

Solid

MOLECULAR WEIGHT:

53.49 g/mol

MELTING POINT:

340°C (640°F);
sublimes at melting
point)

BOILING POINT:

Not applicable

SOLUBILITY:

Soluble in water;
slightly soluble in
alcohol; insoluble in
most organic solvents

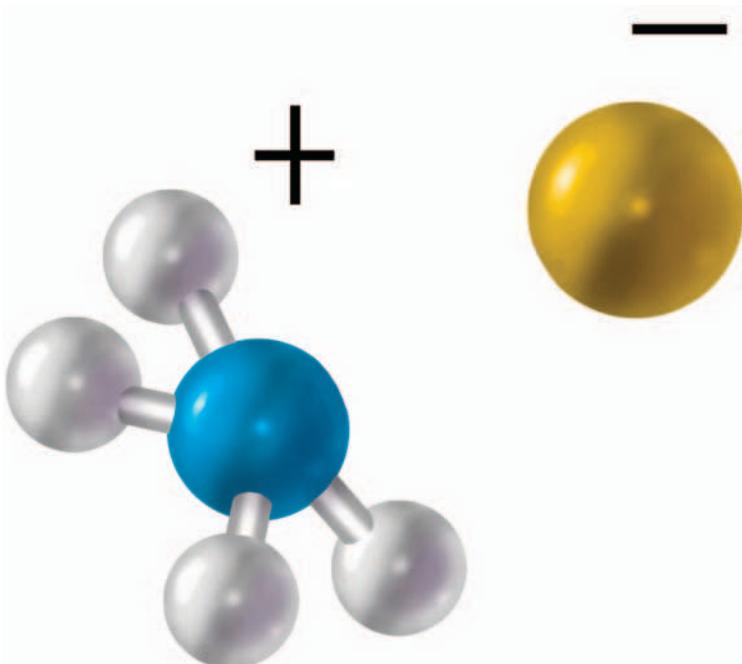
K E Y F A C T S

OVERVIEW

Ammonium chloride (uh-MOH-ni-um KLOR-ide) occurs as odorless white crystals with a cool, salt-like taste. The compound is of interest to historians of chemistry as being one of the first chemicals mentioned by ancient scholars and the first compound of ammonia to have been discovered. For example, the Roman philosopher Pliny the Elder (23-79 CE) wrote about a substance he called hammoniacus sal that appears to have been ammonium chloride. The problem is that various authorities used the term *sal ammoniac* for a variety of materials that were clearly different from each other. No one really knew the compound's actual chemical composition until 1700, when it was discovered by the French botanist Joseph Tournefort (1656-1708). In any case, *sal ammoniac* was an important raw material in early industrial operations, including primarily dyeing and metallurgical operations.

Sal ammoniac is also the name of the mineral form of ammonium chloride. The mineral occurs only rarely in

Ammonium chloride. White atoms are hydrogen; blue atom is nitrogen; and yellow atom is chlorine. PUBLISHERS RESOURCE GROUP



nature, and only then in arid (dry) regions. Since ammonium chloride is quite soluble in water, it remains on the ground only in places where there is little rain. One such location is around the vents of active volcanoes. The compound is formed in these regions when hydrogen chloride (HCl) in volcanic gases and ammonia (NH_3) produced by the decay of plants and animals react with each other to form ammonium chloride, which then settles out onto the ground.

HOW IT IS MADE

One straightforward method of making ammonium chloride is by combining an aqueous solution of some compound of ammonia, usually ammonium sulfate ($(\text{NH}_4)_2\text{SO}_4$), with hydrochloric acid (HCl) and collecting the ammonium chloride that forms by evaporation. Commercially, the compound is obtained as a by-product of the so-called ammonia-soda process for making sodium carbonate (Na_2CO_3). In this process, invented by the Belgian chemist Ernest Solvay (1838-1922) in 1861, ammonia, sodium chloride (NaCl), and carbon dioxide (CO_2) are combined with each other in a series of reactions to make

Interesting Facts

- Sal ammoniac was an important substance in the study of alchemy. The goals and methods of alchemy changed considerably over the period from about the twelfth century to about the sixteenth century and were somewhat different

from culture to culture. For example, one Islamic alchemist, Abu Bakr Mohammed ar-Razi (865–925) classified sal ammoniac as one of the fundamental spirits of matter, along with mercury, sulfur, and arsenic.

sodium carbonate, a very important commercial product. Ammonium chloride is also formed during the reactions and is removed as a by-product.

COMMON USES AND POTENTIAL HAZARDS

Ammonium chloride has a wide variety of commercial uses. One of the best known uses is in dry cell batteries. Dry cell batteries consist of three parts: the anode (the metal bottom of the battery), the cathode (the metal knob at the top of the battery), and the electrolyte (a moist solid material that makes up the body of the battery). Electrons produced in a chemical reaction within the battery flow out of the cathode, through an external circuit (the device to which the battery is attached), back into the battery through the anode, and back to the cathode through the electrolyte. The electrolyte in a dry cell battery consists of a pasty mixture of ammonium chloride with water.

Some other uses of ammonium chloride include:

- As a mordant in dyeing and printing operations;
- As a flux for soldering;
- For the production of other ammonium compounds, especially those used as fertilizers;
- In the manufacture of certain types of polymers, especially the family known as the urea-formaldehyde resins;

Words to Know

ALCHEMY An ancient field of study from which the modern science of chemistry evolved.

AQUEOUS SOLUTION A solution that consists of some material dissolved in water.

ELECTROPLATING Process by which a thin layer of one metal is deposited on top of a second metal by passing an electric current through a solution of the first metal.

FLUX A material that lowers the melting point of another substance or mixture of substances or that is used in cleaning a metal.

MORDANT A substance used in dyeing and printing that reacts chemically with both a dye and the material being dyed to help hold the dye permanently to the material.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

- For the electroplating of metals; and
- As an additive to a licorice-type candy popular in some parts of the world, to which it gives a characteristic salty taste.

Exposure to ammonium chloride may carry certain health risks. When inhaled or deposited on the skin, it may cause irritation of the tissues that may require treatment. The compound is potentially toxic if swallowed. These potential problems are generally of concern primarily to people who work with the compound directly, as in the industries mentioned above. Anyone who is exposed to ammonium chloride should see medical attention immediately.

FOR FURTHER INFORMATION

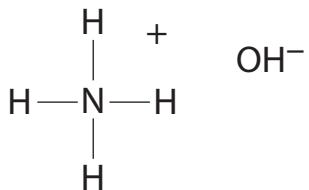
“Ammonium Chloride: Helping to Provide Portable Electricity.”
The Science Center.
http://www.scienceeducation.org/classroom_activities/chlorine_compound/ammonium_chloride.html
(accessed on September 19, 2005).

“Ammonium Chloride, Technical.” Zaclon Incorporated.
http://www.zaclon.com/pdf/amchltec_datasheet.pdf
(accessed on September 19, 2005).

"Material Safety Data Sheet: Ammonium Chloride." Department of Chemistry, Iowa State University.
<http://avogadro.chem.iastate.edu/MSDS/NH4Cl.htm>
(accessed on September 19, 2005).

Patnaik, Praydot. *Handbook of Inorganic Chemicals*. New York: McGraw Hill, 2003.

See Also Ammonia; Ammonium Hydroxide; Sodium Carbonate; Urea



OTHER NAMES:

Ammonia solution,
aqua ammonia,
ammonium hydrate,
spirits of hartshorn

FORMULA:

NH_4OH

ELEMENTS:

Nitrogen, hydrogen,
oxygen

COMPOUND TYPE:

Inorganic base

STATE:

Aqueous solution
(exists only in
solution)

MOLECULAR WEIGHT:

35.05 g/mol

MELTING POINT:

Not applicable

BOILING POINT:

Not applicable

SOLUBILITY:

Not applicable

**K
E
Y
F
A
C
T
S**

OVERVIEW

Ammonium hydroxide (uh-MOH-ni-um hye-DROK-side) is a clear, colorless aqueous solution consisting of ammonia gas dissolved in water. The compound does not exist in any other state. Ammonium hydroxide has a strong pungent, suffocating odor caused by the release of ammonia gas from the solution. Most solutions of ammonium hydroxide range in concentration from less than one percent to about 35 percent ammonia. For most commercial purposes the lowest concentration is about 10 percent ammonia in water.

HOW IT IS MADE

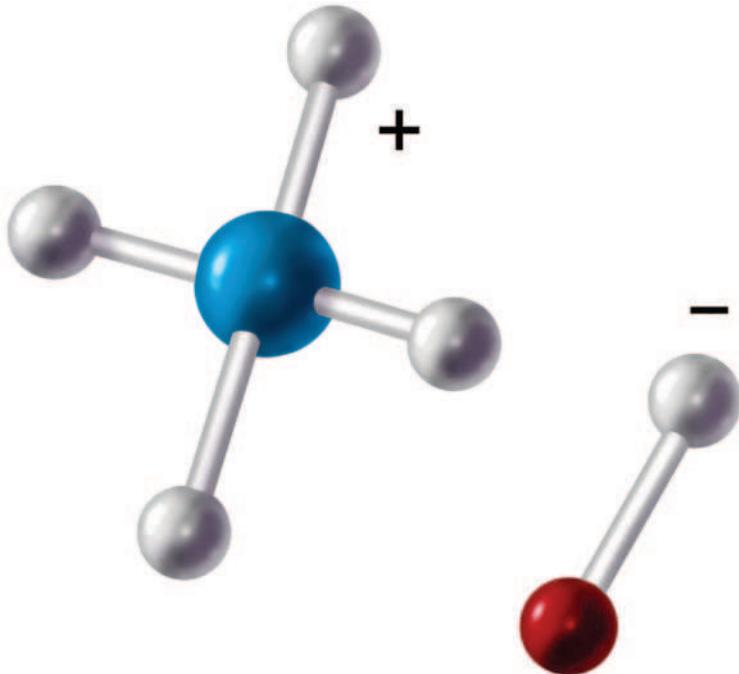
Ammonium hydroxide is prepared by passing ammonia gas (NH_3) into water. Once prepared, ammonium hydroxide solutions tend to be very stable.

COMMON USES AND POTENTIAL HAZARDS

The uses of ammonium hydroxide are closely related to those of ammonia gas from which it is made. The advantage

AMMONIUM HYDROXIDE

Ammonium hydroxide. White atoms are hydrogen; red atom is oxygen; and blue atom is nitrogen. PUBLISHERS RESOURCE GROUP



of using ammonium hydroxide over ammonia is that the reactant (the ammonia) may be more easily controlled when dissolved in water than when available as a gas.

Some uses of ammonium hydroxide include:

- As a cleaning agent in a variety of industrial and commercial products, such as household ammonia, where the concentration of the solution is generally in the range of 3 to 10 percent;
- In the manufacture of rayon and other textiles;
- As a refrigerant;
- As a food additive to maintain the proper level of acidity in the food;
- In the production of soaps and detergents, pharmaceuticals, ceramics, explosives, and inks; and
- In the fireproofing of wood.

Health hazards posed by ammonium hydroxide are a consequence of the ammonia present in the solution. When exposed to the air, such solutions tend to release some ammonia gas, which users may breathe in. The gas may then

Words to Know

AQUEOUS SOLUTION A solution that consists of some material dissolved in water.

PUNGENT Intensely sharp or biting.

cause irritation of the eyes, nose, and throat that, for weak solutions, is generally unpleasant but not dangerous. Exposure to higher concentrations of ammonia released from ammonium hydroxide may result in more serious health problems, such as severe burning of the eyes, nose, and throat and permanent damage to these parts of the body, including blindness, lung disease, and death. One of the most dangerous hazards posed by ammonium hydroxide occurs when it is mixed with substances containing chlorine, such as bleaching products. That combination may result in large amounts of heat and toxic gases that can cause serious injuries and even death.

FOR FURTHER INFORMATION

“Ammonia: Public Health Statement.” Agency for Toxic Substances and Disease Registry.

http://www.atsdr.cdc.gov/toxprofiles/tp126_c1.pdf (accessed on September 19, 2005).

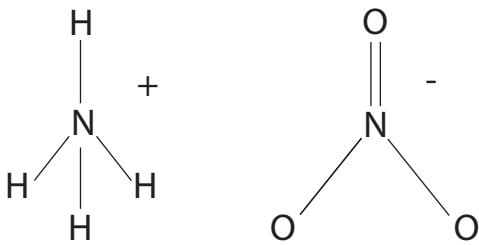
“Ammonium Hydroxide.” Industrial Resources Group, Ltd.

<http://www.indresgroup.com/aqua.htm> (accessed on September 19, 2005).

“Ammonium Hydroxide.” Medline Plus Medical Encyclopedia.

<http://www.nlm.nih.gov/medlineplus/ency/article/002491.htm> (accessed on September 19, 2005).

See Also Ammonia



OTHER NAMES:	German saltpeter; Norway saltpeter; nitric acid, ammonium salt
FORMULA:	NH_4NO_3
ELEMENTS:	Nitrogen, hydrogen, oxygen
COMPOUND TYPE:	Inorganic salt
STATE:	Solid
MOLECULAR WEIGHT:	80.04 g/mol
MELTING POINT:	169.6°C (337.3°F)
BOILING POINT:	211°C (412°F); decomposes at its boiling point
SOLUBILITY:	Very soluble in water, soluble in alcohol and acetone

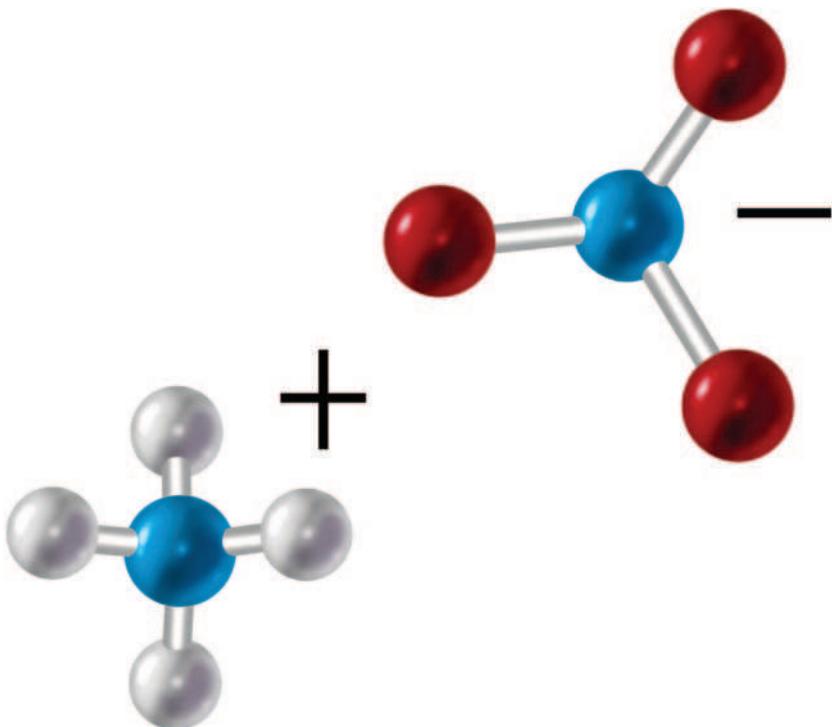
KEY FACTS

OVERVIEW

Ammonium nitrate (uh-MOH-ni-um NYE-trate) is a white crystalline substance first made artificially in 1659 by the German chemist Johann Rudolf Glauber (1604-1670). The compound does not occur in nature because it is so soluble that it is washed out of the soil by rain and surface water. Ammonium nitrate is stable at lower temperatures, but tends to decompose explosively when heated to temperatures above 200°C (390°F). Its two most important uses today are in fertilizers and explosives. In 2004, it ranked fourteenth among all chemicals manufactured in the United States. Just over six million metric tons (6.6 million short tons) of the compound were produced in 2004.

HOW IT IS MADE

Ammonium nitrate is made commercially by passing ammonia gas (NH_3) and a water solution of nitric acid (HNO_3) through a pipe. The ammonia combines with the



Ammonium nitrate. Red atoms are oxygen; white atoms are hydrogen; and blue atoms are nitrogen. PUBLISHERS RESOURCE GROUP

nitric acid to form ammonium nitrate. The formula for this reaction can be written as $\text{NH}_3 + \text{HNO}_3 \rightarrow \text{NH}_4\text{NO}_3$.

Large amounts of heat are released during the reaction, so the pipe and supporting equipment must be very strong. The solution of ammonium nitrate in water is allowed to evaporate, leaving behind pure white crystals of the compound.

COMMON USES AND POTENTIAL HAZARDS

The primary use of ammonium nitrate is the manufacture of fertilizers. In 2005, about 2 million metric tons (2.2 million short tons) of ammonium nitrate fertilizer was used in the United States. The compound is added to soil to provide the nitrogen that plants need to grow. It may be used by itself or in combination with another nitrogen-rich compound, urea, in a mixture known as UAN.

Ammonium nitrate is also an important component of some explosives. It provides the oxygen needed to cause some other material to catch fire and burn very rapidly,

Interesting Facts

- Glauber named the compound he discovered in 1659 *nitrum flammans*, Latin for “flaming nitre.” He chose the name because of the compound’s tendency to explode when exposed to heat.
- An ammonium nitrate explosion in Texas City, Texas, on April 16, 1947, was responsible for the worst industrial accident in U.S. history. While being loaded into two ships at the Texas City harbor, more than 7.5 million kilograms (17 million pounds) of the ammonium nitrate was exposed to flames and exploded. The force of the explosion was so great that it could be felt more than 400 kilometers (250 miles) away in Louisiana. Officials estimate the death toll at 581 people, with more than 5,000 more injured.
- Two American men, Timothy McVeigh and Terry Nichols, used a truckload of ammonium nitrate and other materials to blow up the Alfred P. Murrah Federal Building in Oklahoma City, Oklahoma, on April 19, 1995. The event was one of the worst terrorist incidents ever on American soil.

producing an explosion. One common type of explosive is made of ammonium nitrate mixed with fuel oil and called ANFO (Ammonium Nitrate Fuel Oil). When the mixture is heated, ammonium nitrate breaks down to release oxygen, which causes the rapid combustion (explosion) of the fuel oil.

Some other uses of ammonium nitrate include the following:

- In fireworks, where it provides the oxygen needed to ignite other chemicals;
- In the manufacture of nitrous oxide (N_2O), commonly known as laughing gas;
- In rocket engines, where it provides oxygen needed to burn the rocket fuel;

AMMONIUM NITRATE

- In the manufacture of safety matches, where the compound supplies oxygen to substances that catch fire when the match is struck; and
- As a nutrient in commercial processing for growing yeasts and antibiotics.

FOR FURTHER INFORMATION

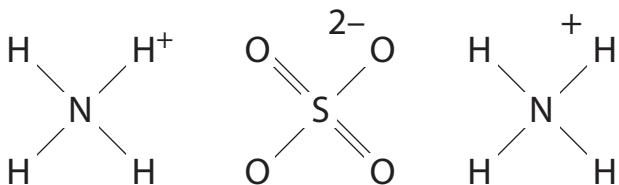
“Chemical Profile for Ammonium Nitrate.” Scorecard: The Pollution Information Site.

http://www.scorecard.org/chemical_profiles/summary.tcl?edf_substance_id=6484522. (accessed on September 13, 2005).

Gorman, Christine. “The Bomb Lurking in the Garden Shed.” *Time* (May 1, 1995): 54.

Olafson, Steve. “The Explosion: 50 Years Later Texas City Remembers.” *The Houston Chronicle* (April 14, 1997): A1. Also available online at
<http://www.chron.com/content/chronicle/metropolitan/txcity/main.html>. (accessed on September 13, 2005).

See Also Ammonia, Nitric Acid, Nitrous Oxide



OTHER NAMES:

Diammonium sulfate;
sulfuric acid,
diammonium salt

FORMULA:

$(\text{NH}_4)_2\text{SO}_4$

ELEMENTS:

Nitrogen, hydrogen,
sulfur, oxygen

COMPOUND TYPE:

Inorganic salt

STATE:

Solid

MOLECULAR WEIGHT:

132.14 g/mol

MELTING POINT:

Not applicable;
decomposes above
 235°C (455°F)

BOILING POINT:

Not applicable

SOLUBILITY:

Soluble in water;
insoluble in alcohol,
acetone, and other
common organic
solvents

K E Y F A C T S

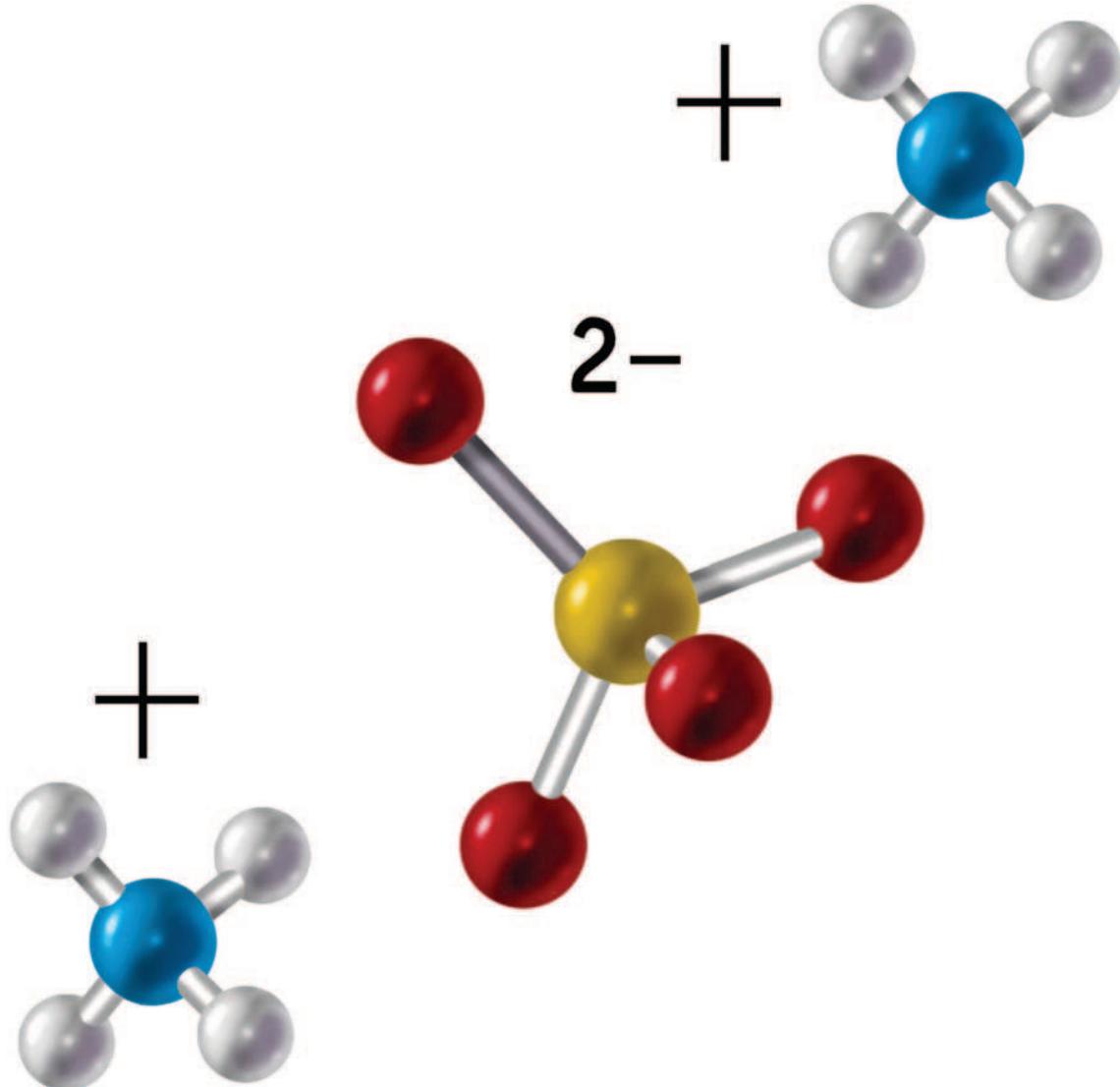
OVERVIEW

Ammonium sulfate (uh-MOH-ni-um SUL-fate) is an odorless, colorless to white crystalline solid that occurs in nature as the mineral mascagnite. In 2004, 2.6 million metric tons (2.9 million short tons) of the compound were produced in the United States, placing it in 21st place among chemicals made in that year. More than 95 percent of the ammonium sulfate produced is used in the production of fertilizers.

HOW IT IS MADE

The primary method of preparation for ammonium sulfate is the direct reaction between ammonia gas (NH_3) and sulfuric acid (H_2SO_4). The ammonium sulfate produced in the reaction is recovered as white crystals after evaporation of the water present in the reaction mixture. Other methods of preparation are also used. For example, gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$) can be treated with ammonia (NH_3) and carbon dioxide (CO_2) to generate ammonium sulfate.

AMMONIUM SULFATE



Ammonium sulfate. Red atoms are oxygen; white atoms are hydrogen; blue atoms are nitrogen; yellow atom is sulfur.
Gray stick indicates double bond. PUBLISHERS RESOURCE GROUP

COMMON USES AND POTENTIAL HAZARDS

Ammonium sulfate is used as a fertilizer because it supplies nitrogen and sulfur, two nutrients that plants need to grow properly. Lesser quantities of the compound are utilized in water treatment plants where it is used to control the acidity of the water being processed. Among other uses of ammonium sulfate are:

Interesting Facts

The German alchemist Andreas Libau (c.1540–1616), better known by his Latinized name of Libavius, wrote what is widely regarded as the first textbook on chemistry in

1597. In that book, *Alchemia*, he described a method for making ammonium sulfate, probably the first mention of the compound in modern sources.

- In the tanning of leather;
- As a fireproofing agent for fabrics and paper;
- In the manufacture of viscose rayon;
- As an additive to cattle feed
- As a nutrient in bacterial cultures;
- In the processing of certain metals, such as chromium and gold; and
- In the manufacture of polymers used in the production of chipboard.

FOR FURTHER INFORMATION

“Ammonium Sulfate.” Hazardous Substances Data Bank.

<http://toxnet.nlm.nih.gov/cgi-bin/sis/search/r?dbs+hsdb:@term+@na+Ammonium+Sulfate> (accessed on September 20, 2005).

Words to Know

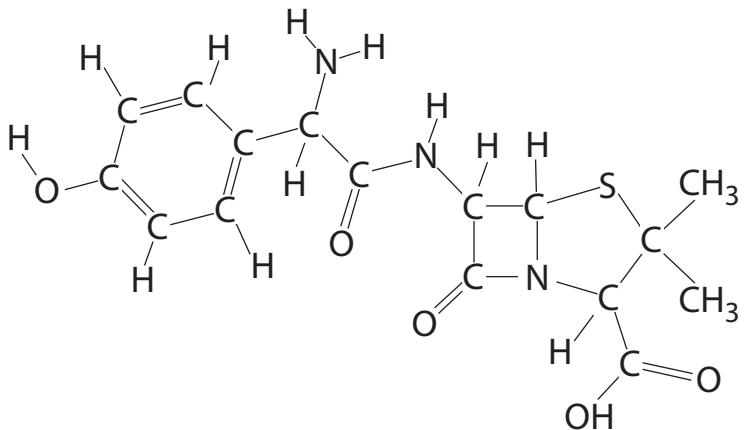
ALCHEMY Ancient field of study from which the modern science of chemistry evolved.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

AMMONIUM SULFATE

- “Ammonium Sulfate Advantage: FAQs.” Honeywell.
<http://www.sulfn.com/main/pages/faqs.asp> (accessed on July 22, 2005).
- “Ammonium Sulfate Industrial Grade.” BASF.
http://www.bASF.de/en/produkte/chemikalien/anorganika/ammonium/ammoniumsulfat_industrial.htm?id=V00 YYuHV7 R07bsf1Gb (accessed on September 20, 2005).
- Bariyanga, Joseph. “Fertilizer.” *Chemistry: Foundations and Applications*. Edited by J. J. Lagowski. Vol. 2. New York: Macmillan Reference USA, 2004.

See Also Ammonia; Sulfuric Acid



OTHER NAMES:

See Overview.

FORMULA:

$C_{16}H_{19}N_3O_5S$

ELEMENTS:

Carbon, hydrogen, nitrogen, oxygen, sulfur

COMPOUND TYPE:

Organic (pharmaceutical)

STATE:

Solid

MOLECULAR WEIGHT:

365.40 g/mol

MELTING POINT:

No data available

BOILING POINT:

No data available

SOLUBILITY:

Soluble in water and alcohol; insoluble in most organic solvents

KEY FACTS

OVERVIEW

Amoxicillin (uh-MOX-uh-sill-in) is also known as D-(-)-alpha-amino-p-hydroxybenzyl penicillin and (2S,5R,6R)-6-[(R) -(-)-2-amino-2-(hydroxyphenyl)acetamido]-3,3-dimethyl-7-oxo-4-thia-1-azabicyclo[3.2.0]heptane-2-carboxylic acid. It is an off-white crystalline antibiotic, a substance that inhibits the growth of microorganisms. It is generally available in the form of the trihydrate (having three molecules of water for each molecule of amoxicillin) with the formula $C_{16}H_{19}N_3O_5S \cdot 3H_2O$. Amoxicillin is effective against a number of disease-causing bacteria, including *streptococci*, the bacterium that cause strep throat, and most strains of *pneumococci*, the bacterium that cause pneumonia, *gonococci*, the bacterium that cause gonorrhea, and *meningococci*, the bacterium that cause meningitis. Amoxicillin works by blocking the formation of bacteria cell walls. Without cell walls, the bacteria die.

Amoxicillin belongs to a group of semisynthetically produced antibiotics called the aminopenicillins. The term



Amoxicillin. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; blue atoms are nitrogen; and green atom is sulfur. Gray sticks are double bonds; striped sticks show a benzene ring. PUBLISHERS RESOURCE GROUP

semisynthetic means that some stages in the preparation of the compound make use of naturally-occurring organisms, such as bacteria or yeast, and some stages involve chemical synthesis. Preparation of the antibiotic penicillin begins with the antibiotic penicillin itself, usually made by one of the molds in the *Penicillium* family, such as *Penicillium chrysogenum* or *Penicillium notatum*. The penicillin molecule is then modified by the addition of a hydroxyl, methyl, or some other group. Other semisynthetic antibiotics related to amoxicillin and penicillin are ampicillin, benzylpenicillin (penicillin G), and phenoxymethylpenicillin (penicillin V).

Scottish bacteriologist Alexander Fleming (1881–1955) discovered penicillin in 1928. The antibiotic was first produced for human use in the early 1940s. Over time, researchers

Interesting Facts

- The original U.S. patent on amoxicillin granted to Beecham Laboratories in 1965 has expired, allowing other pharmaceutical manufacturers to make the drug. This fact accounts for the more than three dozen trade names under which it is sold.

discovered ways of changing the chemical structure of penicillin to create semisynthetic versions. The new antibiotics they created were more effective than penicillin against a wider range of bacteria, often with fewer side-effects. Amoxicillin was discovered by researchers at the Beecham pharmaceutical laboratories in 1962 and marketed about a decade later under the trade name of Amoxil®. The drug is now available under a number of trade names, including Larotid®, Trimox®, Wymox®, Polymox®, and Augmentin®.

HOW IT IS MADE

The preparation of amoxicillin involves a complex series of reactions that begins with penicillin produced by molds or other microorganisms. A variety of chemical reagents is then used to replace one hydrogen atom in the penicillin molecule by the $\text{CH}(\text{NH}_2)\text{C}_6\text{H}_4\text{OH}$ group that converts penicillin into amoxicillin.

COMMON USES AND POTENTIAL HAZARDS

Amoxicillin is available only by prescription and is used against a variety of disease-causing bacteria, such as the cocci bacteria listed above, as well as *Bordetella pertussis*, the bacterium that causes whooping cough, *Salmonella typhi*, the bacterium that causes typhoid, and *Vibrio cholerae*, the bacterium that causes cholera.

The safety of amoxicillin has been extensively tested in trials with both experimental animals and humans. It poses no risk to the vast majority of people. As with all drugs,

Words to Know

ANAPHYLAXIS A serious allergic reaction to a substance that may include respiratory symptoms, fainting, swelling and itching.

REAGENT A substance used in a chemical reaction.

SEMICSYNTHETIC A compound that is prepared in such a way that some stages in

the preparation use naturally-occurring organisms, such as bacteria or yeast, and some stages involve chemical synthesis.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

however, some individuals may experience adverse reactions to the use of amoxicillin. These reactions include fever, skin rash, joint pain, or swollen glands. More serious reactions can include nausea, diarrhea, or anaphylaxis, a potentially fatal allergic reaction. Adverse reactions are most likely to occur among people who have asthma, kidney disease, leukemia, intestinal problems, chronic illness, or a known allergy to other penicillin drugs.

FOR FURTHER INFORMATION

“Amoxicillin.” Drugs.com.

<http://www.drugs.com/amoxicillin.html> (accessed on September 20, 2005).

“Amoxicilin.” Medline Plus.

<http://www.nlm.nih.gov/medlineplus/druginfo/medmaster/a685001.html> (accessed on December 6, 2005).

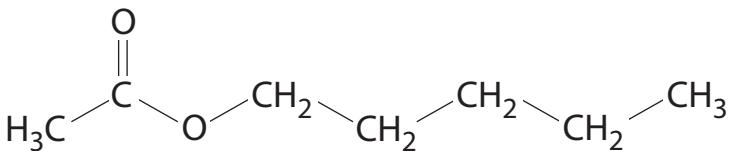
“Amoxicillin.” University of Maryland Medical Center.

<http://www.umm.edu/altmed/ConsDrugs/Amoxicillincd.html> (accessed on September 20, 2005).

“Amoxil® Prescribing Information.” GlaxoSmithKline.

http://us.gsk.com/products/assets/us_amoxil.pdf (accessed on September 20, 2005).

See Also Penicillin



OTHER NAMES:	Pentyl acetate; acetic acid, amyl ester
FORMULA:	$\text{CH}_3\text{COOC}_5\text{H}_{11}$
ELEMENTS:	Carbon, hydrogen, oxygen
COMPOUND TYPE:	Ester (organic)
STATE:	Liquid
MOLECULAR WEIGHT:	130.18 g/mol
MELTING POINT:	-70.8°C (-95.4°F)
BOILING POINT:	149.2°C (300.6°F)
SOLUBILITY:	Slightly soluble in water; soluble in alcohol, ether, and most organic solvents

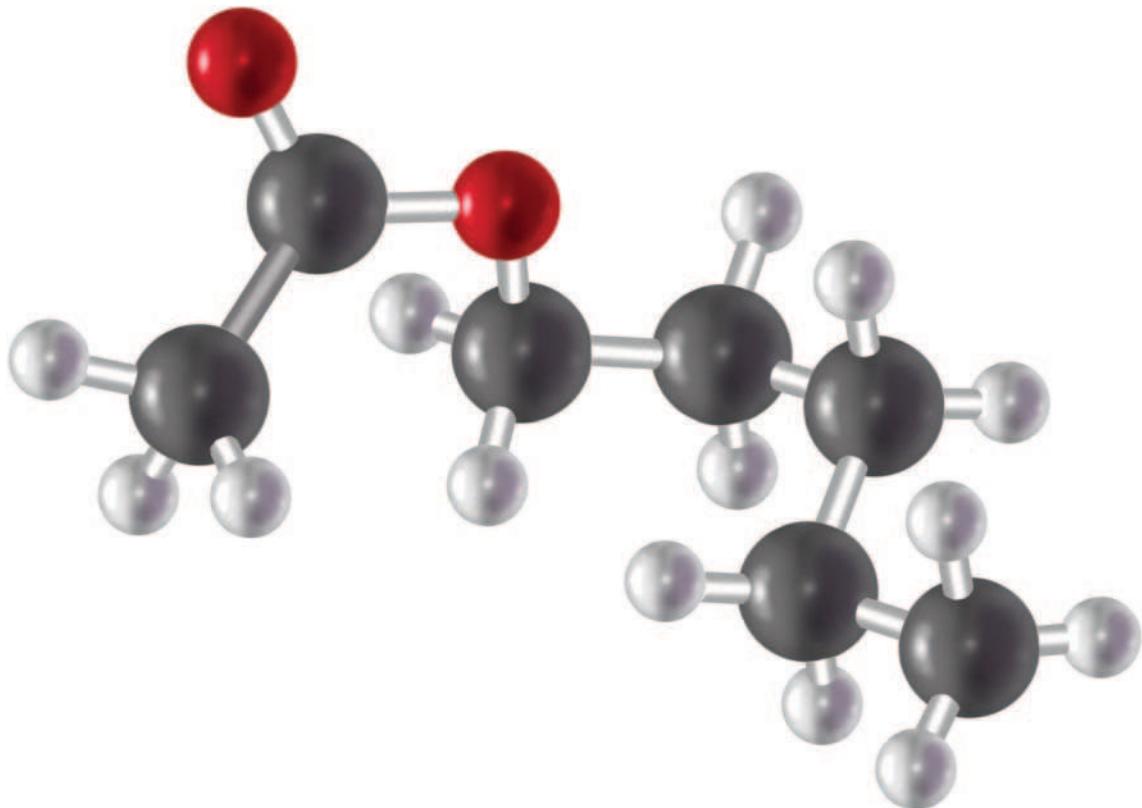
KEY FACTS

OVERVIEW

Amyl acetate (A-mil AS-uh-tate) is a colorless liquid with a distinctive banana-like flavor and odor. Three major isomers of amyl acetate exist: normal (*n*-amyl), secondary (*sec*-amyl), and isoamyl (3-methyl-1-butyl) acetate. Isomers are two or more forms of a chemical compound with the same molecular formula, but different structural formulas and different chemical and physical properties. As an example, the boiling points of the three isomers of amyl acetate are 149.2°C (300.6°F), 142.0°C (287.6°F), and 140.0°C (284.0°F), respectively. Although the amyl acetates are probably best known as flavoring agents because of their distinctive banana-like flavor, they all have a number of interesting industrial applications also.

HOW IT IS MADE

The amyl acetates are made industrially in essentially the same way they are made in a high school or college chemistry



Amyl acetate. Red atoms are oxygen; black atoms are carbon; and white atoms are hydrogen.

Gray stick indicates a double bond. PUBLISHERS RESOURCE GROUP

laboratory. Acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) is added to amyl alcohol ($\text{C}_5\text{H}_{11}\text{OH}$) with a small amount of concentrated sulfuric acid (H_2SO_4) as a catalyst. The specific amyl acetate produced depends on the specific amyl alcohol used in the reaction. The product is separated from the reaction mixture by boiling the liquid mixture.

COMMON USES AND POTENTIAL HAZARDS

Amyl acetate is used as a flavoring agent in the United States and several other countries. It is often blended with other esters to produce flavors that are more fruity and rounded. The compound is also used to flavor products such as chewing gum.

Some applications of amyl acetate depend on its distinctive banana-like odor. The compound may be used, for example, to cover up the unpleasant odors present in certain products,

Interesting Facts

Amyl acetate is the primary component of banana oil and pear oil.

such as shoe polish. Amyl acetate is also used as a warning odor in devices through which gases flow, such as respirators. A warning odor is a distinctive smell that allows users to be aware that a leak exists in a product carrying a gas that has no odor of its own.

Among the many industrial applications of amylic acetate are the following:

- As a solvent for paints and lacquers;
- As an ingredient in fingernail polishes;
- In the production of penicillin; and
- In the manufacture of photographic film, printed and dyed fabrics, and phosphors used in fluorescent lamps.

Amyl acetate is thought to pose a moderate to low health risk. It may cause respiratory problems if inhaled or damage to the skin and eyes if spilled on the body. The most serious safety risk it poses is its flammability. It catches fire easily and burns rapidly, with a moderate risk of exploding under certain circumstances.

FOR FURTHER INFORMATION

“Amyl Acetate.” CAMEO: Conservation and Art Material Encyclopedia Online.

http://www.mfa.org/_cameo/frontend/material_description.asp?name=amyl+acetate&language;=1 (accessed on September 21, 2005).

“n amyl acetate.” CHEMINFO.

<http://intox.org/databank/documents/chemical/amylacet/cie455.htm> (accessed on September 20, 2005).

“n amyl acetate.” New Jersey Department of Health and Senior Services.

<http://www.state.nj.us/health/eoh/rtkweb/1321.pdf> (accessed on September 21, 2005).

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure

ESTER An organic compound formed in the reaction between an organic acid and an alcohol.

ISOMER Two or more forms of a chemical compounds with the same molecular formula, but different structural formulas and different chemical and physical properties

PHOSPHOR A chemical that gives off light when exposed to an electrical current.

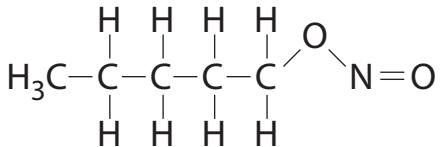
"sec amyl acetate." CHEMINFO.

<http://intox.org/databank/documents/chemical/amylacet/cie468.htm> (accessed on September 20, 2005).

"sec amyl acetate." New Jersey Department of Health and Senior Services.

<http://www.state.nj.us/health/eoh/rtkweb/1643.pdf> (accessed on September 21, 2005).

See Also Acetic Acid; Cellulose Nitrate; Penicillin



OTHER NAMES:

Pentyl nitrite

FORMULA:

$\text{C}_5\text{H}_{11}\text{NO}_2$ or
 $(\text{CH}_3)_2\text{CHCH}_2\text{CH}_2\text{NO}_2$

ELEMENTS:

Carbon, hydrogen,
nitrogen, oxygen

COMPOUND TYPE:

Ester (organic)

STATE:

Liquid

MOLECULAR WEIGHT:

117.15 g/mol

MELTING POINT:

Not available

BOILING POINT:

n-amyl nitrite: 104.5°C
(220°F); isoamyl
nitrite: 99.2°C (210°F)

SOLUBILITY:

Almost insoluble in
water; soluble in
alcohol

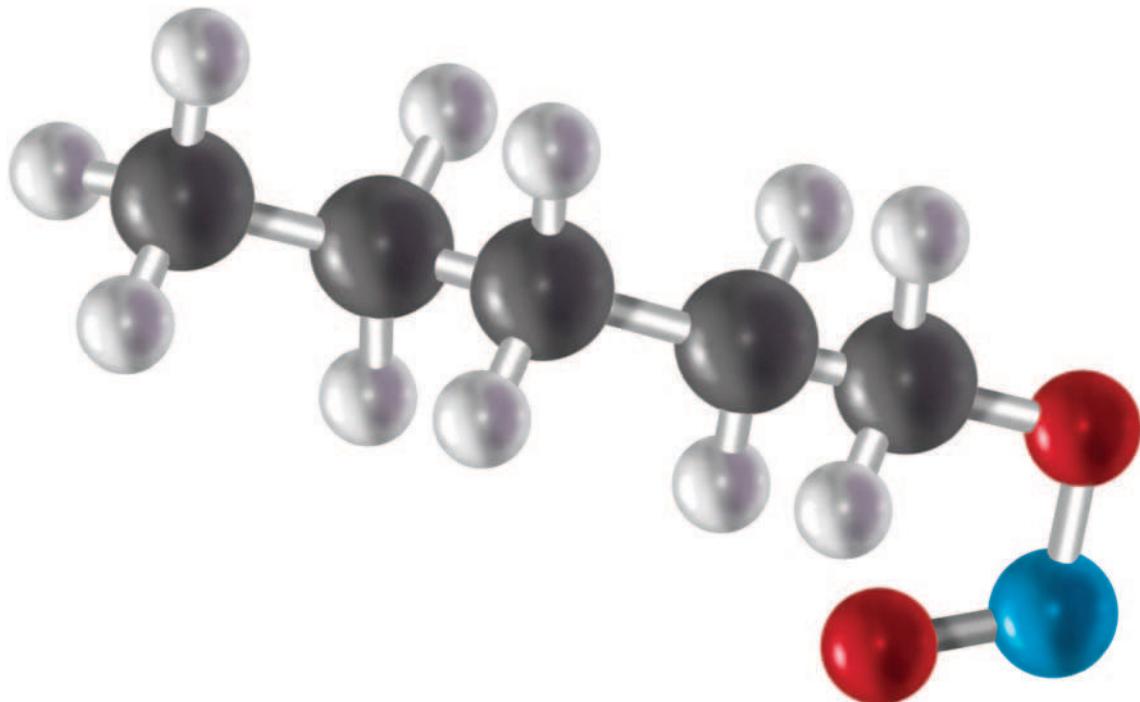
KEY FACTS

OVERVIEW

Amyl nitrite (A-mil NYE-trite) commonly occurs in two isomeric forms: n-amyl nitrite and isoamyl nitrite (3-methylamyl nitrite or 3-methylpentyl nitrite). In common usage, the term amyl nitrite most commonly refers to the isoamyl form. Both isomers occur as yellowish liquids with a distinctive sweet odor and a pungent taste. They are unstable and break down when exposed to air, light, or water. Both isomers are probably best known as vasodilators, substances that cause blood vessels to relax and expand, allowing an increased flow of blood through the body. Because of its widespread use as a hazardous recreational drug, isoamyl nitrite has been banned, except for specified medical uses.

HOW IT IS MADE

Isoamyl nitrite is prepared in a straightforward synthesis in the reaction between isoamyl alcohol ($(\text{CH}_3)_2\text{CHCH}_2\text{CH}_2\text{OH}$) and nitrous acid (HNO_2). The usual method of preparation is



Amyl nitrite. Red atoms are oxygen; black atoms are carbon; white atoms are hydrogen; and the blue atom is nitrogen. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

to combine starch with concentrated nitric acid (HNO_3) to form gaseous nitrous acid. The nitrous acid is then passed through warm isoamyl alcohol. The compound can also be produced by distilling a mixture of potassium nitrite (KNO_2), isoamyl alcohol, and sulfuric acid (H_2SO_4).

COMMON USES AND POTENTIAL HAZARDS

The amyl nitrites are used in medicine for the treatment of pain due to angina, chest pain that occurs because of the constriction of blood vessels. The compounds cause the blood vessels to relax, expand, and ease the flow of blood, with a resulting decrease in chest pain. They are also used by emergency medical providers to treat poisoning by hydrogen cyanide and hydrogen sulfide gas. The amyl nitrites are converted in the body to a compound called methemoglobin, which binds to hydrogen cyanide and hydrogen sulfide, reducing their toxic effects.

U.S. law permits the sale of amyl nitrites as additives to perfumes and for other odorizing purposes. For example, they may be sold as room deodorizers, to eliminate the

Interesting Facts

- Although isoamyl nitrite has legitimate industrial and medical uses, it is probably best known as a recreational drug called “poppers.” The name comes from a popping sound that occurs when a bottle of the liquid is opened. The health consequences of using poppers are such

that the U.S. government passed a ban on the sale of poppers in 1991, except for certain specific purposes. The ban has not eliminated the use of isoamyl nitrite as a recreational drug, since people can still buy the compound for one of its permitted uses.

unpleasant smells that may accumulate in a space. Some people who purchase these room deodorizers or perfumes actually use them as recreational drugs.

An important industrial application of the amyl nitrites is in the production of diazonium compounds. Diazonium compounds contain a characteristic -N=N- grouping and constitute an important family of dyes.

The health risks associated with the use of the amyl nitrites arise because they act as vasodilators. As blood vessels expand and blood flow increases after inhalation of amyl nitrites, blood rushes to the head, and a person may become dizzy and light-headed. These symptoms may provide a “rush”

Words to Know

DISTILLATION A process of separating two or more substances by boiling the mixture of which they are composed and condensing the vapors produced at different temperatures.

ISOMERS Two or more forms of a chemical compound with the same molecular formula,

but different structural formulas and different chemical and physical properties.

VASODILATOR Drug that causes blood vessels to relax and expand.

of excitement that makes them desirable to some people as stimulants. But prolonged use may result in more serious health problems, including headache, nausea, loss of balance, shortness of breath, and rapid heartbeat. For individuals with heart problems, such conditions can lead to heart failure and death.

FOR FURTHER INFORMATION

“Amyl Nitrite.” WholeHealthMD.com.

http://www.wholehealthmd.com/refshelf/drugs_view/1,1524,35,00.html (accessed on December 7, 2005).

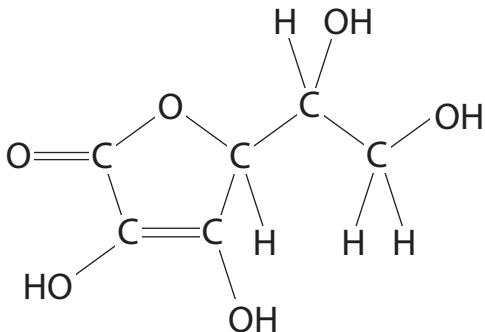
“Amyl Nitrite (Systemic).” Medline Plus.

<http://www.nlm.nih.gov/medlineplus/druginfo/uspdi/202034.html> (accessed on December 7, 2005).

Gahlinger, Paul M. *Illegal Drugs: A Complete Guide to Their History*. Salt Lake City, UT: Sagebrush Press, 2001.

“Isoamyl nitrite.” International Programme on Chemical Safety.

<http://www.inchem.org/documents/icsc/icsc/eics1012.htm> (accessed on December 7, 2005).



OTHER NAMES:

L-ascorbic acid;
vitamin C

FORMULA:

C₆H₈O₆

ELEMENTS:

Carbon, hydrogen,
oxygen

COMPOUND TYPE:

Organic acid

STATE:

Solid

MOLECULAR WEIGHT:

176.12 g/mol

MELTING POINT:

191°C (376°F);
decomposes

BOILING POINT:

Not applicable

SOLUBILITY:

Soluble in water and
alcohol; insoluble in
ether, benzene, and
chloroform

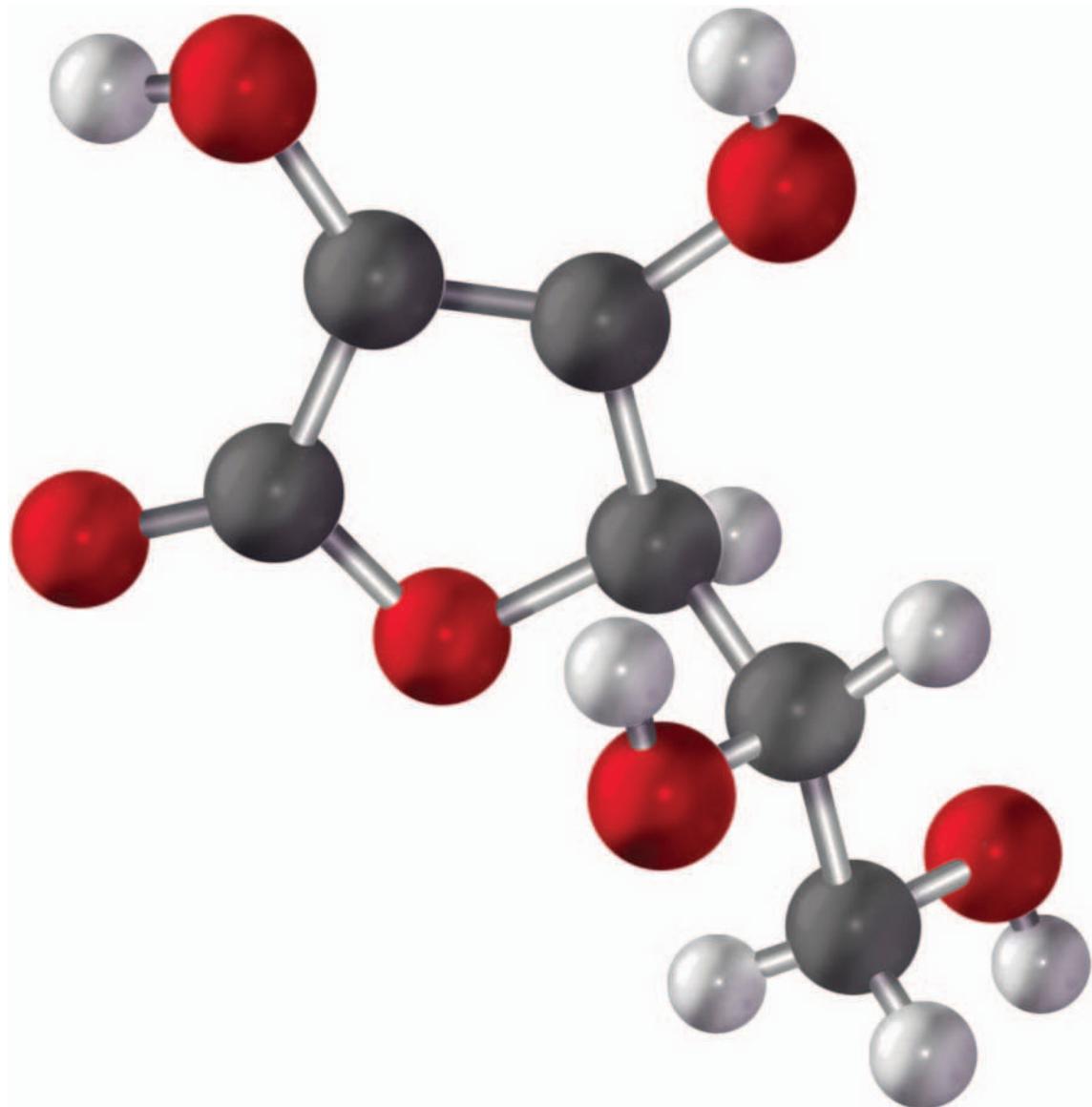
K E Y F A C T S

OVERVIEW

Ascorbic acid (as-KOR-bik AS-id), or vitamin C, is one of the most important dietary vitamins for humans because it plays a crucial role in building collagen, the protein that serves as a support structure for the body. It is a water-soluble vitamin, which means that the body excretes any excess vitamin C in the urine and cannot store a surplus. For that reason, humans must consume vitamin C in their daily diets. Vitamin C is found in many fruits and vegetables and most kinds of fresh meat. Citrus fruits, such as oranges and lemons, are especially rich in the compound.

Humans have known about the consequences of vitamin C deficiency for centuries. People traveling long distances on land or by sea often came down with an illness called scurvy. The same illness struck people living in their own homes during long winters. The disease was characterized by pain and weakness in the joints, fatigue, bleeding gums, tooth loss, slow healing of wounds, and bruising. These symptoms were caused as the body's connective tissue broke down

ASCORBIC ACID



Ascorbic acid. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. Gray sticks are double bonds. PUBLISHERS RESOURCE GROUP

and small blood vessels ruptured. These symptoms began to disappear as fresh foods became more available. If they did not get enough fresh food in their diets, people could die of scurvy.

Scurvy was common enough that many people searched for its cause and cure. Sailors were especially vulnerable to

the disease, and the first recorded investigations involving vitamin C were done by seafaring men. In 1536, French explorer Jacques Cartier (1491–1557) cured his sailors of scurvy by following the advice of Indians in Newfoundland, feeding them extract of pine needles. Scottish physician James Lind (1716–1794) began investigating the disease in 1747. He read many historical accounts of the diseases and combined that information with his own observations to deduce that scurvy occurred only among people with very limited diets. He went on a ten-week sea voyage and fed the soldiers various foods to see which ones were best at curing scurvy. Citrus fruits proved to be most effective in preventing the disease, a result that Lind reported in 1753. Captain James Cook (1728–1779) led expeditions to the South Seas in the late 1700s and kept his crew healthy by feeding them sauerkraut. In 1795 the British navy began serving its sailors a daily portion of lime juice, and two things happened: British sailors stopped getting scurvy, and people began calling sailors “limeys.”

Many people refused to believe that scurvy was caused by a dietary deficiency, suggesting that it was instead the result of eating bad food or lack of exercise. In 1907, Norwegian biochemists Alex Holst (1861–1931) and Theodore Frohlich conducted a study in which guinea pigs were fed an experimental diet that caused them to develop scurvy. The link between the vitamin and the disease was firmly established by this research. Ascorbic acid was first isolated independently by the Hungarian-American biochemist Albert Szent-Györgyi (1893–1986) and the American biochemist Charles Glen King (1896–1988) in 1932. It was synthesized a year later by the English chemist Sir Walter Norman Haworth (1883–1950) and the Polish-Swiss chemist Tadeusz Reichstein (1897–1996), again working independently of each other.

HOW IT IS MADE

Plants and most animals (humans and guinea pigs being two exceptions) synthesize vitamin C in their cells through a series of reactions in which the sugar galactose is eventually converted to ascorbic acid. For many years, the compound has been made commercially by a process known as the Reichstein process, named after its inventor Tadeusz Reichstein. This process begins with ordinary glucose, which is converted to another sugar, sorbitol, which is then fermented to obtain

Interesting Facts

- The vitamin C produced by plants and by synthetic methods are chemically identical and have identical effects in the human body.
- American chemist Linus Pauling (1901–1994) believed that very large doses of vitamin C could prevent and cure the common cold and

other flu-like diseases. Because of his reputation (he won two Nobel Prizes, one in chemistry, as well as the Peace Prize), his viewpoint was highly respected. He was not, however, able to prove his theory to the satisfaction of his scientific colleagues.

yet another sugar, sorbose. The sorbose is then converted step-by-step into a series of other products, the last of which is ascorbic acid.

Chemists have long been searching for an alternative to the Reichstein process because it uses so much energy and produces by-products that are hazardous to the environment. In the 1960s, Chinese scientists developed a method that involves only two steps in the synthesis of ascorbic acid, and in the early 2000s, Scottish scientists were attempting to develop a method that involved only a single step using fermentation. Currently, however, the Reichstein process remains the most popular method for making the compound.

COMMON USES AND POTENTIAL HAZARDS

The best known use of vitamin C is as a nutritional supplement, taken to ensure that one receives his or her daily minimum requirement of the vitamin. The recommended daily allowance (RDA) of vitamin C for adults is 60 milligrams per day. Anyone who eats a well-balanced diet that includes citrus fruits, tomatoes, and green leafy vegetables probably does not need to take a vitamin supplement. However, the amount of vitamin C one normally receives from a supplement is unlikely to cause any harm.

Words to Know

REDUCTION Chemical reaction in which oxygen is removed from a substance or electrons are added to a substance.

SYNTHESIS Chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

In addition to its nutritional uses, ascorbic acid has a number of other industrial applications, including:

- As a food preservative;
- As a reducing agent in chemical processes;
- As a preservative in foods;
- As a color fixing agent in meats, helping meats keep their bright red appearance;
- As an additive to bread dough, where it helps increase the activity of yeast used in the dough; and
- As a treatment for abscission in citrus plants, the tendency for a plant to lose its leaves, flowers, and fruits.

FOR FURTHER INFORMATION

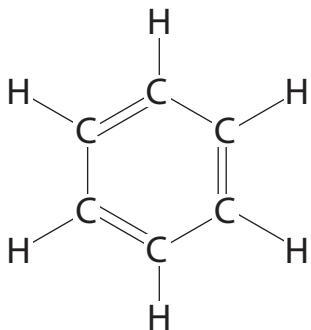
“Ascorbic Acid.” International Programme on Chemical Safety.
<http://www.inchem.org/documents/icsc/icsc/eics0379.htm>
 (accessed on September 21, 2005).

Carpenter, Kenneth J. *The History of Scurvy and Vitamin C*. Cambridge, UK: Cambridge University Press, 1986.

“L Ascorbic Acid.” University of Texas at Austin College of Engineering Department of Biomedical Engineering.
<http://www.engr.utexas.edu/bme/ugrad/UGLab/resources/MSDS/ascorbic%20acid.pdf> (accessed on September 21, 2005).

Mead, Clifford, and Thomas Hager, eds. *Linus Pauling: Scientist and Peacemaker*. Portland, OR: Oregon State University Press, 2001.

See Also Collagen; Glucose



OTHER NAMES:	Benzol; cyclohexatriene
FORMULA:	C_6H_6
ELEMENTS:	Carbon, hydrogen
COMPOUND TYPE:	Aromatic hydrocarbon (organic)
STATE:	Liquid
MOLECULAR WEIGHT:	78.11 g/mol
MELTING POINT:	5.49°C (41.9°F)
BOILING POINT:	80.09°C (176.2°F)
SOLUBILITY:	Slightly soluble in water; soluble in alcohol, ether, and acetone

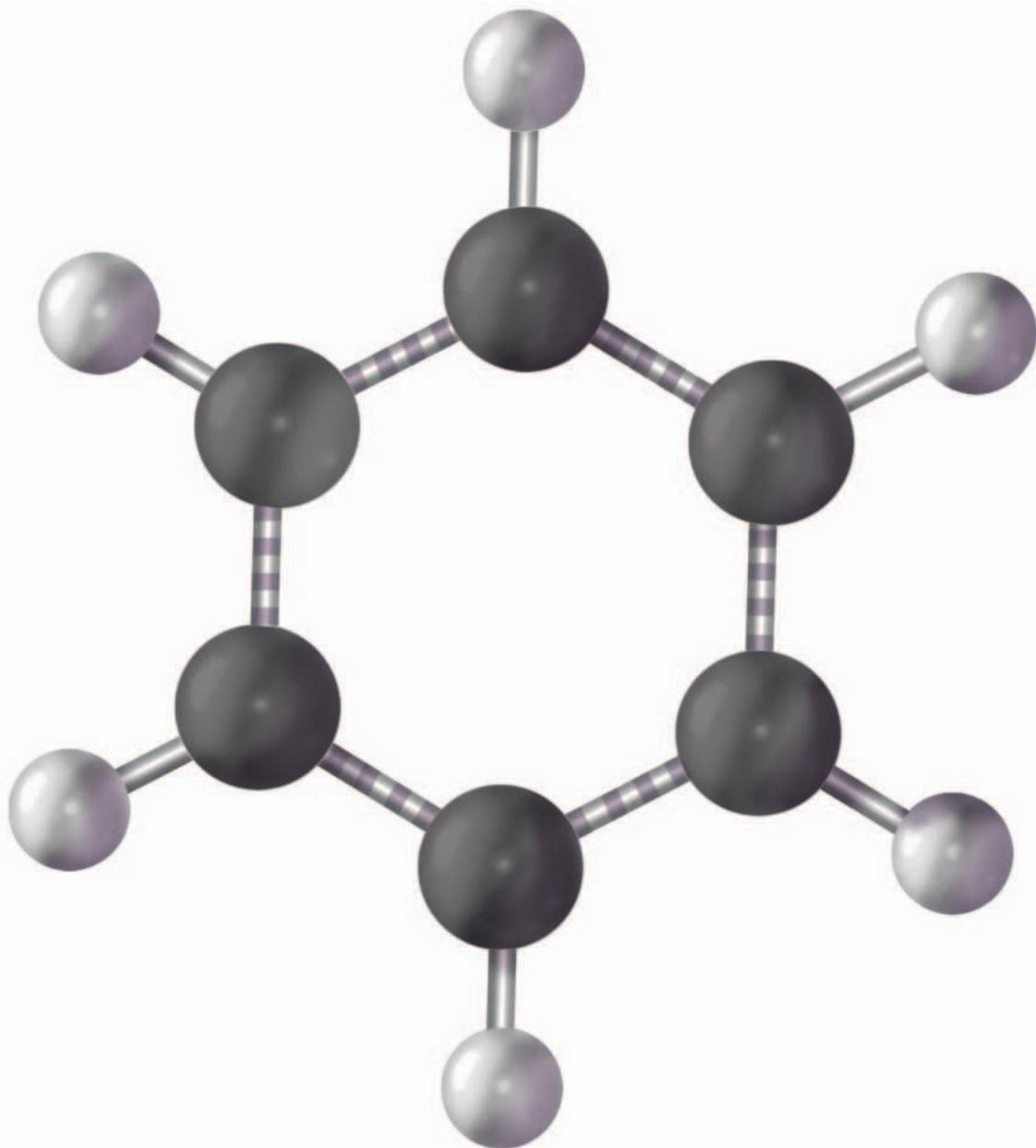
KEY FACTS

OVERVIEW

Benzene (BEN-zeen) is a clear, colorless liquid with an aromatic (fragrant) odor. It occurs in coal and petroleum, from which it is extracted for commercial use. Benzene is very flammable, burning with a smoking flame. The compound was discovered in 1825 by the English chemist and physicist Michael Faraday (1791–1867), who gave the compound the name of bicarburet of hydrogen. It was given its modern name of benzene (benzin, at the time) by the German chemist Eilhardt Mitscherlich (1794–1863).

The chemical structure of benzene remained one of the great mysteries in chemistry for nearly half a century. The compound's formula, C_6H_6 , suggests that it contains three double bonds. A double bond consists of four electrons that hold two atoms in close proximity to each other in a molecule. Yet benzene has none of the chemical properties common to double-bonded substances. The solution to this problem was suggested in 1865 by the German chemist Friedrich August Kekulé (1829–1896). Kekulé suggested that the six

BENZENE



Benzene. White atoms are hydrogen and black atoms are carbon. Striped sticks indicate a benzene ring. PUBLISHERS RESOURCE GROUP

carbon atoms in the benzene molecule are arranged in a ring, with one hydrogen atom attached to each carbon. The ring itself consists of three double bonds and three single bonds, alternating with each other in the ring. The fact that

the double bonds in benzene do not act like double bonds in other compounds was explained by the German chemist Johannes Thiele (1860–1935), who suggested that the bonds in benzene shift back and forth between single and double bonds so rapidly that they are not able to behave like typical double bonds. Chemists now use a variety of chemical formulas for representing the character of chemical bonds in benzene.

Benzene is a very popular raw material for a variety of industrial chemical reactions. In 2004, U.S. manufacturers produced 8.8 million metric tons (9.7 million short tons) of benzene, placing it in twelfth place among all chemicals made in the United States that year.

HOW IT IS MADE

At one time, benzene was obtained from coal tar, the thick gooey liquid left over after soft coal is converted to coke. This method has now been largely replaced by a variety of methods that use crude oil or refined petroleum as a raw material. In the most popular of these methods, toluene ($C_6H_5CH_3$) from petroleum is heated over a catalyst of platinum metal and aluminum oxide (Al_2O_3). The toluene loses its methyl group ($-CH_3$), leaving benzene as the primary product. Other methods are available for changing the molecular structure of hydrocarbons found in petroleum and converting them to benzene.

COMMON USES AND POTENTIAL HAZARDS

By far the most important use of benzene is as a raw material in the synthesis of other organic compounds. More than 90 percent of the benzene produced in the United States is used to make ethylbenzene (55 percent), cumene (24 percent), and cyclohexane (12 percent). The first two compounds rank fifteenth and twentieth, respectively, among all chemicals produced in the United States each year. Another five percent of benzene production goes to the synthesis of a large variety of other organic compounds, including nitrobenzene, chlorobenzene, and maleic anhydride, a raw material for the manufacture of plastics. Smaller amounts of benzene are used as a solvent for

Interesting Facts

- Kekulé's discovery of the formula for benzene is one of the most interesting in the history of chemistry. The story is told that he worked so hard on the problem that he often dreamed about the compound at night. One evening, he dreamed

of a snake with a tail in its mouth. Kekulé immediately awoke, went to his work table, and drew a structure for the benzene molecule inspired by the snake: a molecule in the shape of a ring made of carbon atoms.

cleaning purposes, in chemical reactions, and as a gasoline additive.

As with most chemicals, benzene can enter the body in one of three ways: through the skin, the nose, or the throat. People who handle or work with benzene in their workplaces are at greatest risk of exposure to benzene and should take precautions in working with the material. Because of its serious health hazards, benzene is no longer included in most materials with which the average person comes into contact. On those occasions when a person does come into contact with benzene, first aid and medical attention should be sought for treatment of the exposure.

The health effects of exposure to liquid benzene or benzene fumes depends on the amount of benzene taken into the body. The most common symptoms of benzene exposure include irritation of the mucous membranes, convulsions, depression, and restlessness. At greater doses, a person may experience respiratory failure, followed by death. Even at low concentrations, benzene can cause long-term effects for people who are regularly in contact with the compound. The most important of these effects are carcinogenic. Benzene is known to cause damage to bone marrow, resulting in a form of cancer of the blood known as leukemia.

Words to Know

CARCINOGEN A chemical that causes cancer in humans or other animals.

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

MUCOUS MEMBRANES Tissues that line the moist inner lining of the digestive,

respiratory, urinary, and reproductive systems.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

FOR FURTHER INFORMATION

“Chronic Toxicity Summary: Benzene.” California Office of Health Hazard Assessment.
http://www.oehha.org/air/chronic_rels/pdf/71432.pdf
(accessed on September 21, 2005).

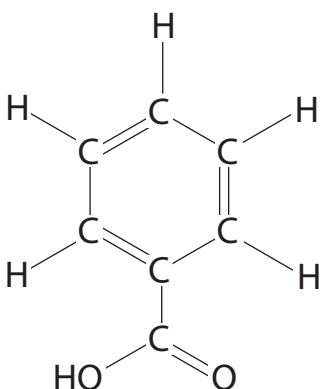
Newton, David E. “Benzene.” In *Gale Encyclopedia of Science*. Edited by K. Lee Lerner and Brenda Wilmoth Lerner. 3rd ed., vol. 1. Detroit: Gale, 2004.

“Spectrum Chemical Fact Sheet: Benzene.” Spectrum Laboratories.
<http://www.speclab.com/compound/c71432.htm> (accessed on September 21, 2005).

“Toxicity Summary for Benzene.” The Risk Assessment Information System.
<http://risk.lsd.ornl.gov/tox/profiles/benzene.shtml> (accessed on September 21, 2005).

“Toxicological Profile for Benzene.” Agency for Toxic Substances and Disease Registry.
<http://www.atsdr.cdc.gov/toxprofiles/tp3.html> (accessed on September 21, 2005).

See Also Benzoic Acid; Styrene

**OTHER NAMES:**

Carboxybenzene,
benzenecarboxylic
acid, phenylformic
acid

FORMULA:

C_6H_5COOH

ELEMENTS:

Carbon, hydrogen,
oxygen

COMPOUND TYPE:

Organic acid

STATE:

Solid

MOLECULAR WEIGHT:

122.12 g/mol

MELTING POINT:

$122.35^\circ C$ ($252.23^\circ F$)

BOILING POINT:

$249.2^\circ C$ ($480.6^\circ F$)

SOLUBILITY:

Slightly soluble in
water; soluble in
alcohol, ether, and
acetone

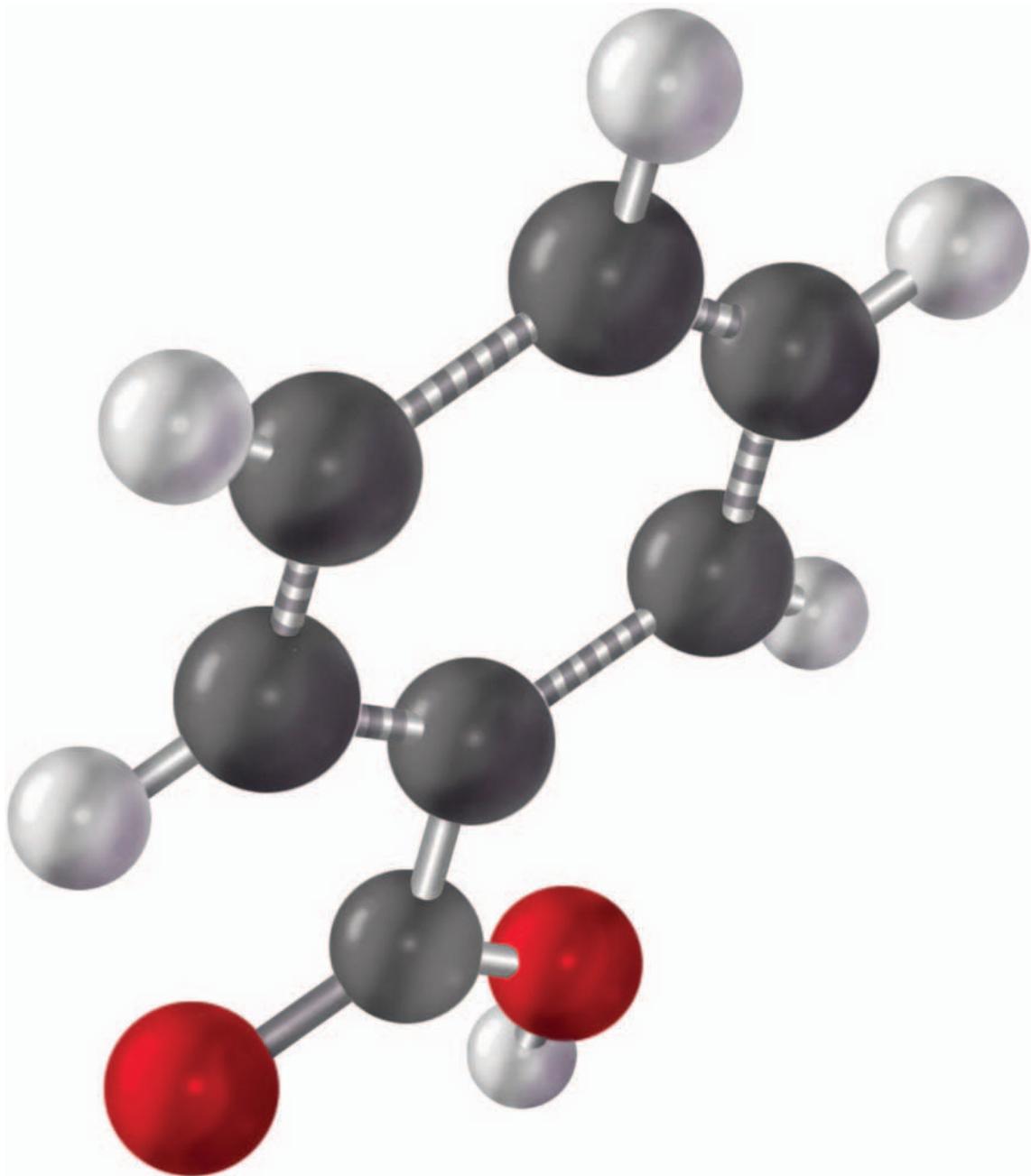
KEY FACTS**OVERVIEW**

Benzoic acid (ben-ZO-ik AS-id) is the simplest of the aromatic carboxylic acids, a family of organic compounds containing the carboxyl (-COOH) group. It occurs in the form of white crystalline needles or thin plates. Many naturally occurring plants contain benzoic acid, including most types of berries and the natural product called gum benzoin, a plant common to the islands of Java, Sumatra, and Borneo. Gum benzoin may contain up to 20 percent benzoic acid. Benzoic acid is also excreted by most animals (except fowl) in the form of a related compound called hippuric acid ($C_6H_5CONHCH_2COOH$).

HOW IT IS MADE

Some benzoic acid is prepared from gum benzoin and other natural compounds containing high concentrations of the compound. Three methods are available for making benzoic acid commercially. In one method, toluene ($C_6H_5CH_3$) is

BENZOIC ACID



Benzoic acid. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon.

PUBLISHERS RESOURCE GROUP

oxidized at high temperature and pressure over a cobalt catalyst. In a second procedure, phthalic anhydride ($C_6H_4(CO)_2O$) is heated to remove carbon dioxide, with the formation of benzoic acid. In the third method, toluene is

first converted to benzotrichloride ($C_6H_5Cl_3$), which is then hydrolyzed to obtain benzoic acid.

COMMON USES AND POTENTIAL HAZARDS

More than half of all the benzoic acid produced in the United States is used in the manufacture of various polymeric products, primarily the family of plastics known as the polyvinyl acetates (PVAs). The PVAs, in turn, are used as adhesives, caulk, sealants, and coatings for paper, film, and plastic foil. About a quarter of all benzoic acid is converted to its sodium and potassium salts, sodium benzoate (C_6H_5COONa) and potassium benzoate (C_6H_5COOK), for use as food preservatives. Sodium benzoate and potassium benzoate are now the most widely used food preservatives in the world. They are added to a host of products, such as soft drinks and fruit juices, jams and jellies, baked goods, and salad dressings. They are also added to a number of non-food products such as mouthwashes, toothpastes, cosmetic creams, and deodorants.

Other uses of benzoic acid include:

- As an additive to antifreezes, where it acts to prevent corrosion of engine parts;
- As a seasoning agent in the processing of tobacco products;
- As a stabilizing agent in the processing of photographs; and
- For the treatment of fungal infections.

Benzoic acid and its sodium and potassium salts pose moderate health hazards to people who work directly with them. They may cause skin, nose, and eye problems if inhaled or deposited on the body. In the form of finely divided dust, they may also pose a fire or explosive risk. Most people do not encounter any of these compounds in a form in which they pose a health risk. All three compounds are considered safe by the U.S. Food and Drug Administration as food additives provided their concentration in foods does not exceed 0.1 percent. Other nations have set a safe level for the three compounds as food additives as high as 1.25 percent.

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

HYDROLYSIS The process by which a compound reacts with water to form two new compounds.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

FOR FURTHER INFORMATION

"Benzoic Acid." Chemical Land 21.

<http://www.chemicalland21.com/arokorhi/industrialchem/organic/BENZOIC%20ACID.htm> (accessed on September 22, 2005).

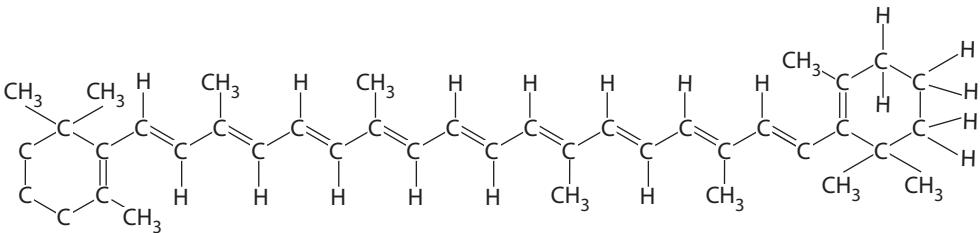
"Benzoic Acid and Sodium Benzoate." International Programme on Chemical Safety.

<http://www.inchem.org/documents/cicads/cicads/cicad26.htm> (accessed on September 22, 2005).

"Benzoic Acid (CASRN 65 85 0)." U.S. Environmental Protection Agency.

<http://www.epa.gov/iris/subst/0355.htm> (accessed on September 22, 2005).

See Also Denatonium Benzoate

**OTHER NAMES:**

β,β -Carotene;
carotaben

FORMULA:

$C_{40}H_{56}$

ELEMENTS:

Carbon, hydrogen

COMPOUND TYPE:

Hydrocarbon
(organic)

STATE:

Solid

MOLECULAR WEIGHT:

536.87 g/mol

MELTING POINT:

183°C (361°F)

BOILING POINT:

Not applicable

SOLUBILITY:

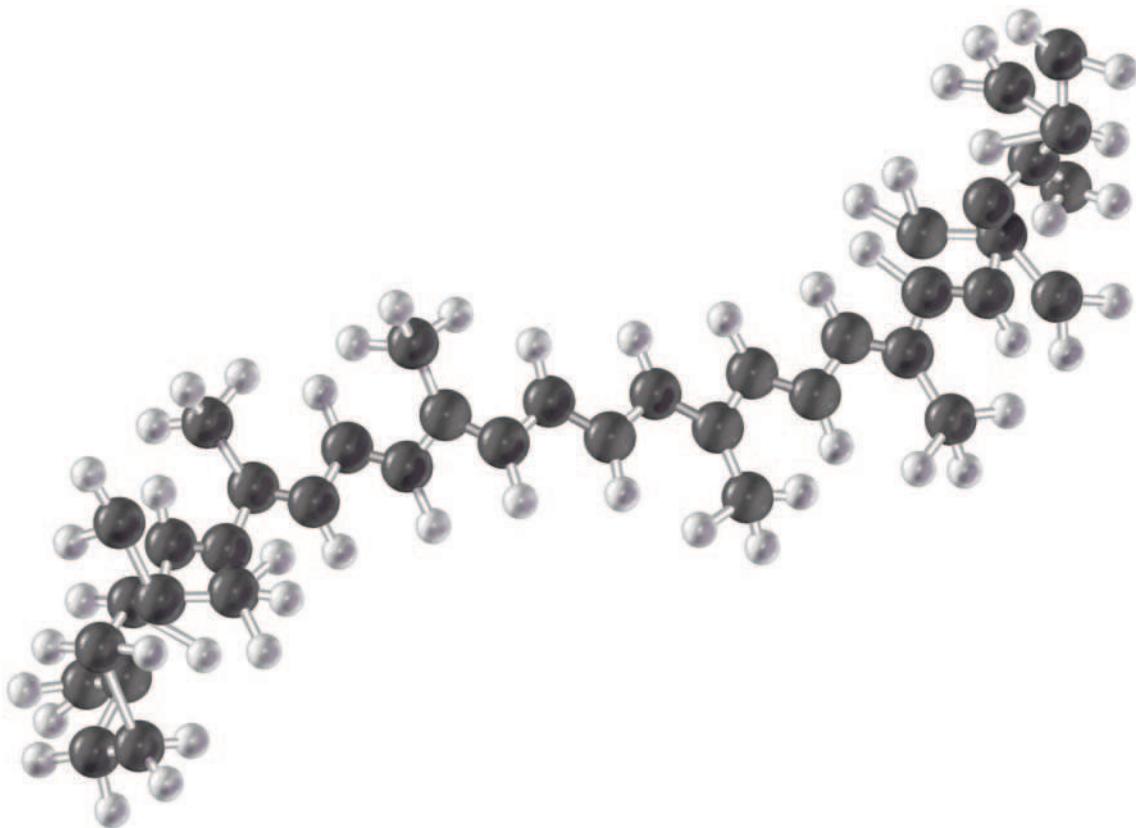
Insoluble in water;
slightly soluble in
alcohol; soluble in
ether, acetone,
benzene, and fats

KEY FACTS**OVERVIEW**

Beta-carotene (β -carotene; BAY-tuh KARE-oh-teen) belongs to a family of organic compounds called the carotenoids. The carotenoids are all brightly pigmented (colored) compounds found in a number of plants, bacteria, algae, and fungi. Beta-carotene is responsible for the yellowish to orange color of pumpkins, apricots, sweet potatoes, nectarines, and, most notably, carrots. The compound also occurs in spinach and broccoli, but in such small concentrations that the green chlorophyll present masks the orange color of beta-carotene. In its pure form, beta-carotene occurs as purple crystals shaped like thin leaflets.

In plants, algae, and photosynthetic bacteria, beta-carotene plays an important role in photosynthesis, the process by which plants convert water and carbon dioxide into carbohydrates and oxygen. In nonphotosynthetic bacteria and fungi, beta-carotene protects the organism against the harmful effects of light and oxygen.

Animals require beta-carotene for normal growth and development, but are unable to manufacture the compound



Beta carotene. White atoms are hydrogen and black atoms are carbon. PUBLISHERS RESOURCE GROUP

themselves. As a result, they must ingest some beta-carotene from plant sources in order to stay healthy. The compound is a provitamin, a substance that is converted in the body to a vitamin. Beta-carotene is converted into vitamin A, whose role in the body is the maintenance of strong bones and teeth and healthy skin and hair.

Beta-carotene also acts as an antioxidant, a substance that attacks free radicals in the body that may cause cancer. It may also protect against heart disease and strengthen the body's immune system.

Beta-carotene was first isolated by the German chemist Heinrich Wilhelm Ferdinand Wackenroder (1789–1854), who extracted the compound from carrot roots in 1831. The compound was first synthesized in 1950 by the Swiss chemist Paul Karrer (1889–1971).

Interesting Facts

- The name carotene comes from the Latin word for “carrot.”
- Scientists have found a way to change the genetic structure of rice so that it contains beta-carotene. The modified rice is designed to help people in countries where vitamin A deficiency is a serious problem.
- At least eight other carotenoids are known. They include α -carotene, β,ψ -carotene, and ψ,ψ -carotene.

HOW IT IS MADE

Beta-carotene can be obtained from natural sources by crushing or pulverizing the source (such as carrots) and adding a solvent that will dissolve the organic components of the plant. These components can then be separated from each other by chromatographic techniques. A major commercial source of beta-carotene obtained by this method is the algae *Dunaliella salina*, which grows in large salt lakes in Australia. The compound can also be prepared synthetically by one of two methods, the BASF and the Roche methods, both named after the pharmaceutical firms where they were developed. Both methods of preparation begin with long-chain hydrocarbons containing about twenty carbon atoms each. These hydrocarbons are then joined to each other to form the 40-carbon beta-carotene compound.

COMMON USES AND POTENTIAL HAZARDS

Beta-carotene has two uses: in vitamin supplements and as a food additive. Anyone who eats a healthy diet that includes foods rich in vitamin A, such as fish oil, liver, eggs, butter, and orange or yellow vegetables and fruits, will get adequate amounts of beta-carotene. However, many

Words to Know

ANTIOXIDANT A chemical that attacks free radicals, chemical structures that attack cells and may be responsible for the development of cancer.

CHROMATOGRAPHY A process by which a mixture of substances passes through a column consisting of some material that causes the individual components in the mixture to separate from each other.

FREE RADICAL An atom or group of atoms with a single unpaired electron. Free

radicals are very active chemically and tend to attack and destroy other compounds. They are responsible for damage in cells that may lead to a number of health problems, including cancer and ageing.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

people take vitamin supplements to ensure that they have enough beta-carotene (as well as other vitamins) in their daily diet. Although some warnings have been issued about taking too much vitamin A, there is no clinical evidence that an overdose of the vitamin does any long-term harm to a person.

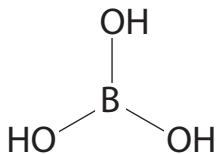
Beta-carotene is used as a food additive to increase the color intensity of a product. It is used primarily with yellow and orange foods, such as butter and margarine, although it is sometimes added to ice cream and fruit juices as well. Beta-carotene is used in only very small amounts as a food additive. In these amounts, it poses no health hazard to humans or other animals. The compound has also been used in experiments to test its effectiveness against certain diseases, such as lung cancer. In such cases, it has been found to be more harmful than beneficial, increasing the risk of cancer and death among people participating in the studies.

FOR FURTHER INFORMATION

“Beta carotene.” University of Bristol School of Chemistry.
http://www.chm.bris.ac.uk/motm/carotene/beta_carotene_home.html (accessed on September 22, 2005).

- “Beta carotene.” University of Maryland Medical Center.
<http://www.umm.edu/altmed/ConsSupplements/BetaCaroteneCS.html> (accessed on September 22, 2005).
- Palvetz, Barry A. “A Bowl of Hope, Bucket of Hype?” *The Scientist* (April 2, 2001): 15.
- “Taking Supplements of the Antioxidant.” *Consumer Reports* (September 2003): 49.

See Also Chlorophyll



OTHER NAMES:

Orthoboric acid;
hydrogen orthoborate;
boracic acid

FORMULA:

H_3BO_3

ELEMENTS:

Hydrogen, boron,
oxygen

COMPOUND TYPE:

Inorganic acid

STATE:

Solid

MOLECULAR WEIGHT:

61.83 g/mol

MELTING POINT:

170.9°C (339.6°F)

BOILING POINT:

Decomposes above
its melting point

SOLUBILITY:

Somewhat soluble in
water, ethyl alcohol,
and glycerol

K
E
Y
F
A
C
T
S

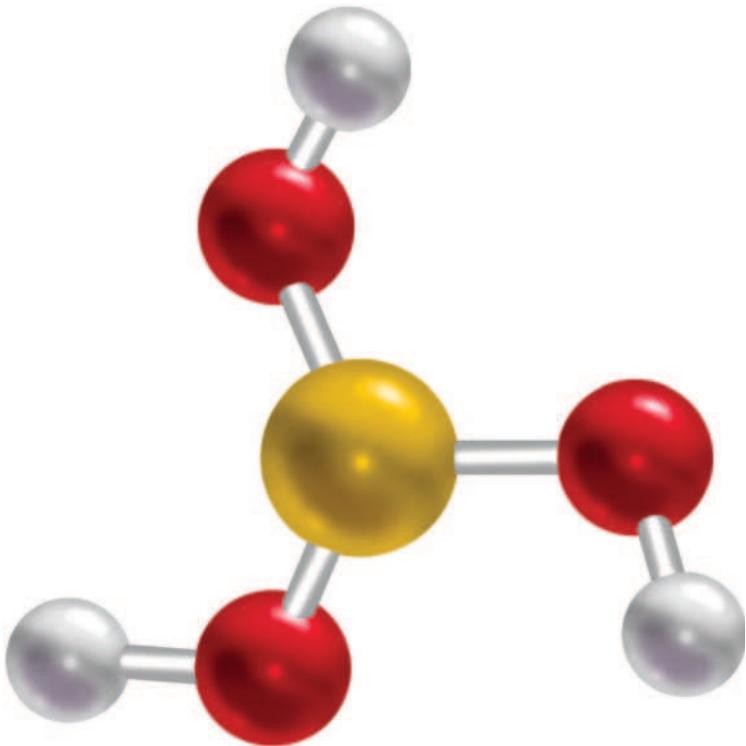
Boric Acid

OVERVIEW

Boric acid (BORE-ik ASS-id) is a colorless, odorless, white or colorless powder or crystalline material with a slightly oily feeling that slowly decomposes with heat, changing first to metaboric acid (HBO_2), then to pyroboric acid ($\text{H}_2\text{B}_4\text{O}_7$), and eventually to boric oxide (B_2O_3). The compound's solubility is very much a factor of temperature. In cold water, about 5 grams (0.2 ounce) of boric acid dissolve in 100 mL (3.4 ounces) of water, while at 100°C (212°F), its solubility increases to 25 grams (0.9 ounce) in 100 mL (3.4 ounces) of water.

Boric acid occurs naturally in a number of locations where it has precipitated out of hot springs. It may occur then in the form of the mineral sassolite. Given its abundance in nature, it is not surprising that the compound has been known to and used by humans for many centuries. For example, the Greeks are known to have used boric acid as an antiseptic, to preserve foods, and as a cleaning agent. The first person to prepare boric acid in Europe is thought to be

Hydrogen orthoborate. Red atoms are oxygen; white atoms are hydrogen; and yellow atom is boron. PUBLISHERS RESOURCE GROUP



the German chemist Wilhelm Homberg (1652-1715), who in 1702 treated natural borax ($\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$) with acid to obtain a product he called sedative salt, probably a form of boric acid. The chemical structure of the compound was determined in 1808 by the French researchers Louis-Joseph Gay-Lussac (1778-1850) and Louis-Jacques Thénard (1777-1857).

HOW IT IS MADE

The most common method of producing boric acid is by treating the relatively abundant borax with hydrochloric or sulfuric acid and crystallizing out the boric acid that forms in the reaction. A less common method of preparation involves the treatment of borax brine solutions with a chelating agent that binds to the borates present in the brine, which can then be converted to boric acid.

Interesting Facts

- The most common sources of boric acid are hot springs and volcanic sites such as the mineral springs at Vichy and Aix-la-Chapelle in France and Wiesbaden in Germany, around the volcanic regions of Tuscany in Italy, and in dry lakes of California and Nevada, such as California's Borax Lake.
- Boric acid was once used as an ingredient in talcum and diaper powders and in salves for diaper rashes because of its antiseptic properties. The compound was eventually banned for such uses, however, when regulators decided that it was too toxic if it accidentally entered the body through an open wound or by being swallowed.

COMMON USES AND POTENTIAL HAZARDS

By far the greatest amount of boric acid is used in the production of heat-resistant (borosilicate) glass, glass fibers, porcelain enamels, crockery, laboratory glassware, and other specialized types of glass and ceramic materials. In 2004, about 80 percent of all the boric acid used in the United States was applied to these purposes. Although accounting for a much smaller amount of boric acid, another well-known use of the compound is as an antiseptic in eyewashes, ointments, and mouthwashes. It is also used as a preservative in foods and as a fungicide on citrus fruit crops. In the United States, a residue of no more than 8 parts per million is permitted on fruits treated with boric acid.

Some other important applications of boric acid include the following:

- In the production of flame retardant materials;
- As a pesticide for the treatment of cockroaches, black carpet beetles, termites, fleas, fire ants, centipedes, grasshoppers, and slugs;
- As a flux in welding and brazing;
- In the synthesis of many boron compounds;

Words to Know

BRINE SOLUTION A solution that is saturated with or nearly saturated with sodium chloride or other inorganic salts.

CHELATING AGENT An organic compound that binds (“grabs on to”) some specific

compound or compounds present in a mixture.

FLUX A material that lowers the melting point of another substance or mixture of substances or that is used in cleaning a metal.

- To provide the finishing touches on the production of leather and fur products;
- In the manufacture of latex paints; and
- In the nickel plating of metallic products.

Boric acid is toxic if swallowed. It causes nausea, vomiting, diarrhea, and stomach cramps. In extreme cases, it can cause the collapse of the circulatory system, delirium, convulsions, coma, and death. Ingestion of no more than 5 grams (0.2 ounce) of boric acid can cause death in an infant. Ingestion of 15 to 20 grams (0.5 to 0.7 ounce) by an adult can also be fatal. Boric acid can also cause irritation of the skin, which, in extreme cases, can result in a condition known as borism. Borism is characterized by dry skin, eruptions of the skin and mucous membranes, and gastric disturbances.

FOR FURTHER INFORMATION

“Borax; Boric Acid, and Borates.” IPM of Alaska.

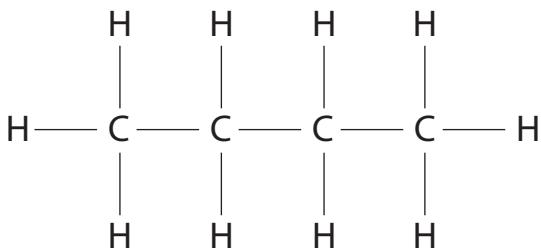
<http://www.ipmofalaska.com/files/Borates.html> (accessed on October 12, 2005).

“Boric Acid: Technical Description.” Manufacturas Los Andes.

http://www.mandes.com.ar/technic_boric_acid.php (accessed on October 12, 2005).

Potter, Mike. “Cockroach Elimination.” University of Kentucky Entomology.

<http://www.uky.edu/Agriculture/Entomology/entfacts/struct/ef614.htm> (accessed on October 12, 2005).



OTHER NAMES:	n-Butane
FORMULA:	C_4H_{10}
ELEMENTS:	Carbon, hydrogen
COMPOUND TYPE:	Alkane; saturated hydrocarbon (organic)
STATE:	Gas
MOLECULAR WEIGHT:	58.12 g/mol
MELTING POINT:	-138.3°C (-216.9°F)
BOILING POINT:	-0.5°C (31°F)
SOLUBILITY:	Slightly soluble in water; soluble in ethyl alcohol, ether, and chloroform

KEY FACTS

Butane

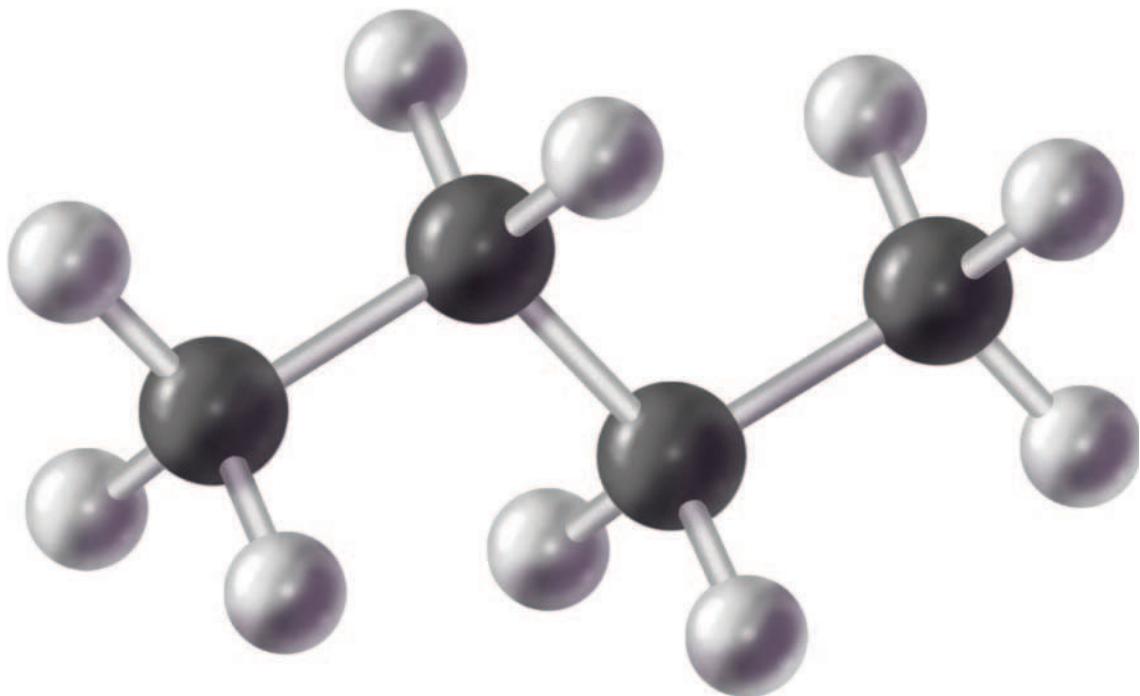
OVERVIEW

Butane (BYOO-tane) is a colorless gas with the odor of natural gas that is highly flammable and explosive. It exists in two isomeric forms. Isomers are forms of a chemical compound with the same molecular formula (in this case, C_4H_{10}), but different structural arrangements. In one isomer ("normal" or "n-" butane), the four carbon atoms are arranged in a continuous chain, while in the other ("iso-butane"), three carbon atoms are arranged in a continuous chain and the fourth carbon atom is attached to the middle atom in that chain.

Butane occurs naturally in natural gas, where it is present to the extent of about 1 percent, and in petroleum, where it exists only in very small amounts. Butane is used primarily as a fuel and as a chemical intermediary, a compound used to produce other chemical substances.

HOW IT IS MADE

Butane is obtained during the separation of natural gas into its components. Natural gas consists primarily (70 to



Butane. Black atoms are carbon; white atoms are hydrogen. All sticks are single bonds. PUBLISHERS RESOURCE GROUP

90 percent) of methane. Other components include ethane (about 9 percent), propane (about 3 percent), and butane (about 1 percent). The remaining 1 to 2 percent of natural gas is impurities, including nitrogen, carbon dioxide, compounds of sulfur, and water. The process of separating natural gas into its components begins with the removal of water. The natural gas is forced through some drying agent, such as diethylene glycol ($\text{HO}(\text{CH}_2)_2\text{O}(\text{CH}_2)_2\text{OH}$), which efficiently removes water from the gas. The dry gas is then treated with an agent to remove carbon dioxide, compounds of sulfur, and other impurities. Diethanolamine ($\text{HO}(\text{CH}_2)_2\text{NH}(\text{CH}_2)_2\text{OH}$) is an efficient “scavenger” of many of these impurities.

After dehydration and removal of impurities, the remaining gas consists almost entirely of simple hydrocarbons, primarily methane, ethane, propane, and butane. These hydrocarbons can be separated from each other by cooling them until they reach the point at which they become liquid. As each component liquefies, it can be removed from the gas and further purified. For example, butane changes from a gas to a liquid at -0.5°C (31°F) and propane liquefies at about -42°C (-44°F). So as the

Interesting Facts

- Some people have used butane as a recreational drug, inhaling it to get “high.” The practice is extremely dangerous, however, and has led some users to die of asphyxiation (suffocation).
- Researchers at Surrey Satellite Technology, Ltd. in the United Kingdom have built a small satellite the size of a soccer ball that is propelled by pressurized butane gas. The butane thruster was built for \$15,000 and was launched aboard a Russian Cosmos spacecraft in June 2000.
- Contact with liquid butane or LPG can cause frostbite, for which first aid involves washing with cold water.

hydrocarbon mixture cools, liquid butane can be removed before any other components liquefy. In actual practice, liquid butane and propane are removed together, leaving behind only ethane and methane. The mixture of liquid butane and propane is known as liquid petroleum gas (LPG). Since propane and butane have many similar properties, there is often no commercial reason to go to the expense of separating them from each other. LPG is typically stored above or below ground in large, insulated containers and shipped in insulated trucks or rail cars.

COMMON USES AND POTENTIAL HAZARDS

Butane, in both gaseous and liquid form, is widely used as a fuel. Because it is easily stored and transported in small containers, it is used as a fuel for backpacking stoves, small space heaters, and portable torches. Butane is also the fuel most commonly used in cigarette lighters. It may be used in its pure form or in combination with propane as liquefied petroleum gas (LPG). Butane is also used as a fuel for larger household and industrial operations.

In terms of volume, one of the major uses of butane is in the synthesis of other organic compounds, especially synthetic rubber and high-octane liquid fuels used in aviation. The gas

Words to Know

EUPHORIA A state of extreme happiness and enhanced well-being.

HALLUCINATIONS Visions or other perceptions of things that are not really present.

is sometimes added to gasoline in cold climates to improve the rate at which the fuel evaporates and burns. A relatively new application for butane is as a propellant for spray products, such as hair spray and spray paints, and as a refrigerant. Butane is being used for these purposes as a substitute for chlorofluorocarbons (CFCs), which have been shown to have damaging effects on the Earth's ozone layer. Small amounts of butane are also used as food additives, usually in foods that are dispensed as sprays.

The primary health hazard posed by butane is its narcotic effects. When inhaled, either accidentally or intentionally, it produces a sequence of bodily changes that include, at first, a sense of euphoria and excitement. Increased doses may then produce harmful results, such as nausea, vomiting, sneezing, coughing, blurred vision, slurred speech, and increased salivation. Even higher doses result in confusion, perceptual distortion, hallucinations, and delusions. Eventually, a sequence of life-threatening conditions may develop, including depression of the central nervous system, irregular heartbeat, drowsiness, coma, and death.

FOR FURTHER INFORMATION

"Acute Exposure Guideline Levels (AEGLs) for n butane." National Advisory Council, U.S. Environmental Protection Agency. Washington, DC: U.S. Environmental Protection Agency, April 2004. Also available online at <http://www.epa.gov/oppt/aegl/pubs/tsd102.pdf> (accessed on December 29, 2005).

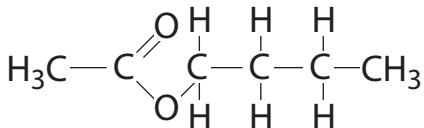
"Butane." International Labour Organization. http://www.ilo.org/public/english/protection/safework/cis/products/icsc/dtasht_icsc02/icsc0232.htm (accessed on December 29, 2005).

"Butane." International Programme on Chemical Safety.

<http://www.inchem.org/documents/pims/chemical/pim945.htm#SectionTitle:2.1%20%20Main%20risks%20and%20target%20organs> (accessed on December 29, 2005).

Russell, Justin. "Fuel of the Forgotten Deaths." *New Scientist* (February 6, 1993): 21–23.

See Also Methane; Propane

**OTHER NAMES:**

Three forms exist;
see Overview

FORMULA:

C₆H₁₂O₂; see Overview
for expanded
formulas

ELEMENTS:

Carbon, hydrogen,
oxygen

COMPOUND TYPE:

Carboxylic acids
(organic)

STATE:

Liquid

MOLECULAR WEIGHT:

116.16 g/mol

MELTING POINT:

-98.9°C (-146°F) to
-78°C (-108°F)

BOILING POINT:

95.1°C (203°F) to
126.1°C (259.0°F)

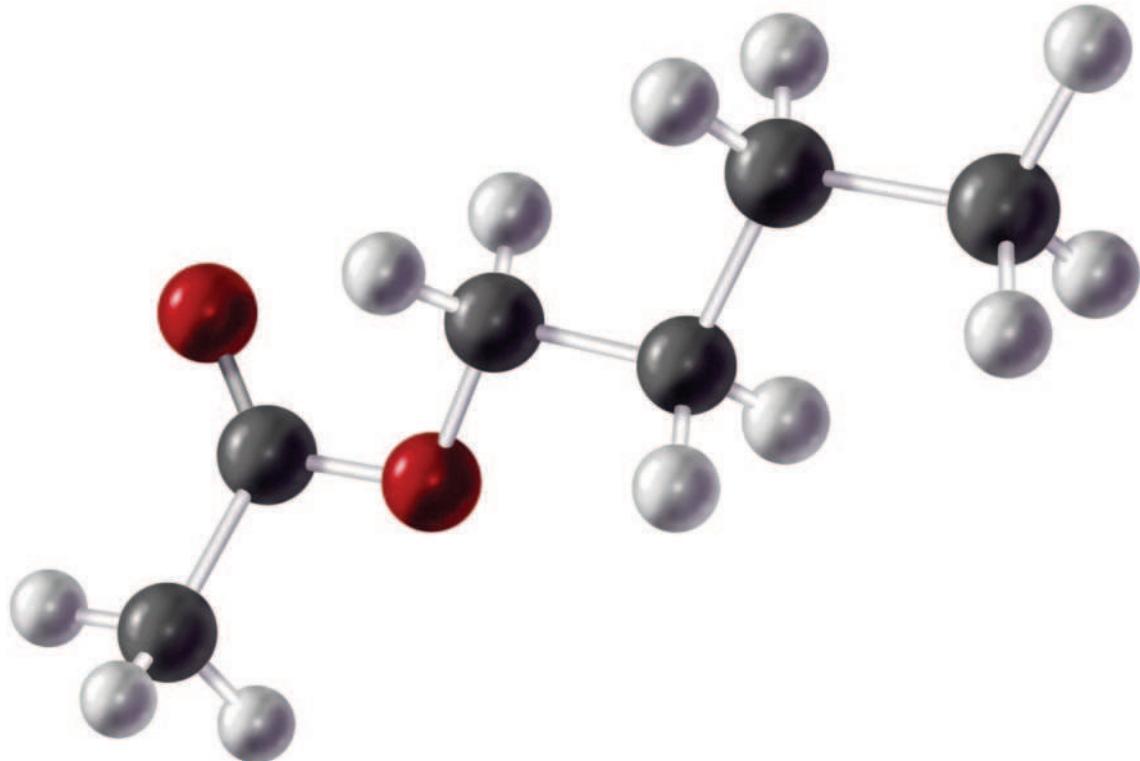
SOLUBILITY:

Insoluble or slightly
soluble in water;
soluble in alcohol and
ether

K E Y F A C T S**OVERVIEW**

Butyl acetate (BYOO-til AS-uh-tate) exists in three isomeric forms. Isomers are two or more forms of a chemical compound with the same molecular formula, but different structural formulas and different chemical and physical properties. Their names and synonyms are as follows: *n*-butyl acetate is known as butyl ester, or butyl ethanoate; *sec*-butyl acetate is called 1-methylpropyl ester and acetic acid *sec*-butyl ester; and *tert*-butyl acetate is known by 1,1-dimethylethyl ester or acetic acid, *tert*-butyl ester. The expanded chemical formulas for the forms are slightly different: *n*-butyl acetate, CH₃COOCH₂CH₂CH₂CH₃; *sec*-butyl acetate, CH₃COOCH(CH₃)(C₂H₅); and *tert*-butyl acetate, CH₃COOC(CH₃)₃.

About 100,000 kilograms (250,000 pounds) of the *n*-butyl isomer, the most popular, is produced each year in the United States. By far the most common use of the compound (and the isomers) is as an industrial solvent, with about 90 percent of total output going for this use.



Butyl acetate. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

HOW IT IS MADE

All three isomers of butyl acetate are made by reacting acetic acid (CH_3COOH) with the appropriate butyl alcohol (*n*-butyl alcohol, *sec*-butyl alcohol, or *tert*-butyl alcohol, respectively).

COMMON USES AND POTENTIAL HAZARDS

The primary use of all three isomers of butyl acetate is as a solvent in the production of lacquers and paints; photographic film; resins; and coatings for furniture, fixtures, containers, and automobiles. Since the 1990s, these compounds have been substituted for other solvents that are considered hazardous environmental pollutants.

Other uses of the isomers are as follows:

- *n*-butyl acetate: as an odor enhancer in perfumes; in the synthesis of synthetic flavors, pharmaceuticals, and other organic compounds;

Interesting Facts

- Scientists have discovered that the butyl acetates are present in the so-called alarm pheromones secreted by animals, particularly insects. An alarm pheromone is a chemical secreted when

an animal is under attack. It alerts other animals in the neighborhood to the attack and “tells” them to respond to the attack that has been detected.

- sec*-butyl acetate: as a thinner for nail enamels and a solvent for leather finishes;
- tert*-butyl acetate: as a gasoline additive.

Relatively little data are available on the health effects of exposure to the isomers of butyl acetate. The chemical is known to cause skin rash, respiratory problems, and burning of the eyes upon contact. And by comparing it with other similar organic compounds, the assumption is made that long-term exposure to high concentrations would result in damage to the nervous system. Since most people do not encounter butyl acetate in the pure form, these risks are minimal. Only workers in one of the fields in which the chemical is used need take special precautions to use care in working with the amyl acetates.

Words to Know

ISOMER One of two or more forms of a chemical compound with the same molecular formula, but different structural formulas and different chemical and physical properties.

SOLVENT A substance that is able to dissolve one or more other substances

FOR FURTHER INFORMATION

"Global Ethyl, Butyl Acetate Demand Expected to Rebound." *The Oil and Gas Journal* (April 24, 2000): 27.

"n butyl acetate." New Jersey Department of Health and Senior Services.

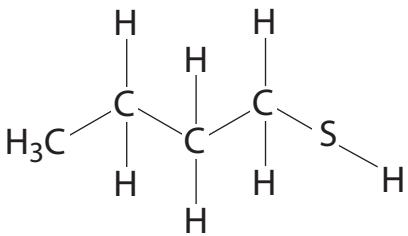
<http://www.state.nj.us/health/eoh/rtkweb/1329.pdf> (accessed on September 22, 2005).

"Occupational Safety and Health Guideline for tert butyl Acetate." Occupational Health and Safety Administration.

<http://www.osha.gov/SLTC/healthguidelines/tertbutylacetate/recognition.html> (accessed on September 22, 2005).

"sec butyl acetate." Scorecard: The Pollution Information Site.

http://www.scorecard.org/chemical_profiles/summary.tcl?edf_substance_id=105%2d46%2d4 (accessed on September 22, 2005).



OTHER NAMES:

Three forms of butyl mercaptan exist; see Overview

FORMULA:

C₄H₁₀S

ELEMENTS:

Carbon, hydrogen
sulfur

COMPOUND TYPE:

Mercaptan (organic)

STATE:

Liquid

MOLECULAR WEIGHT:

90.19 g/mol

MELTING POINT:

Forms vary greatly;
see Overview

BOILING POINT:

Forms vary; see
Overview

SOLUBILITY:

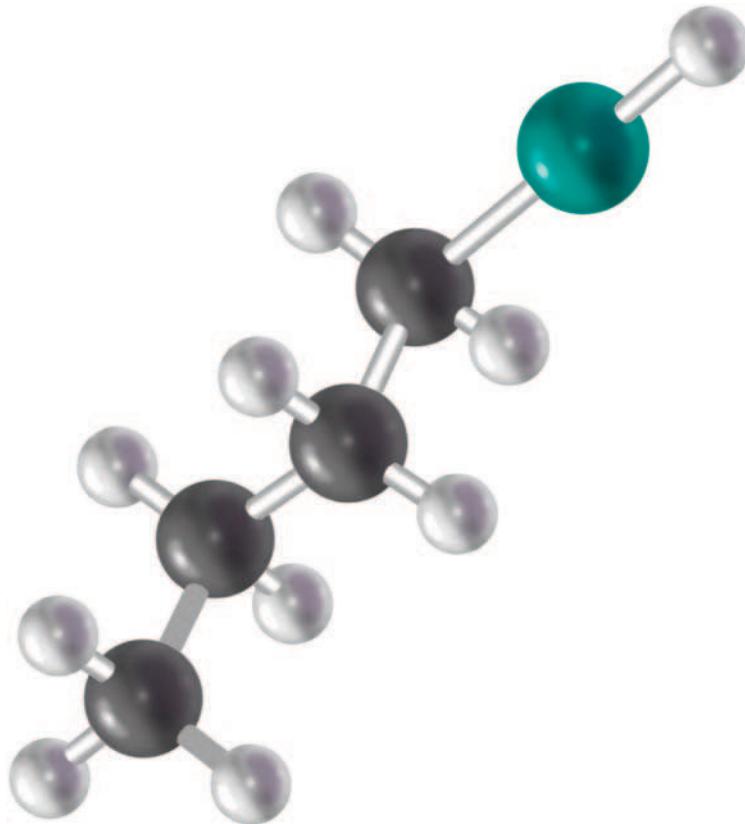
Insoluble or slightly
soluble in water
(depending on form);
soluble in alcohol and
ether

K E Y F A C T S

OVERVIEW

Three isomers of butyl mercaptan (BYOO-til mer-KAP-tan) exist. Isomers are two or more forms of a chemical compound with the same molecular formula, but different structural formulas and different chemical and physical properties. The names and synonyms of the isomers are: *n*-butyl mercaptan, known as 1-butanethiol, 1-mercaptopbutane, or *n*-butyl thioalcohol; *sec*-butyl mercaptan, called 2-butanethiol or 1-methyl-1-propanethiol; and *tert*-butyl mercaptan, also known as 2-methyl-2-propanethiol. The isomers have slightly different expanded chemical formulas: *n*-butyl mercaptan, CH₃CH₂CH₂CH₂SH; *sec*-butyl mercaptan, CH₃CH₂CH(SH)CH₃; and *tert*-butyl mercaptan, CH₃C(CH₃)(SH)CH₃. The isomers have vastly differing melting points: *n*-butyl mercaptan, -115.67°C (-176.21°F); *sec*-butyl mercaptan, -165°C (-265°F); and *tert*-butyl mercaptan, -0.5°C (31°F). Boiling points differ, but not as greatly: *n*-butyl mercaptan, 98.46°C (209.2°F); *sec*-butyl mercaptan, 85°C (180°F); and *tert*-butyl mercaptan, 64°C (150°F).

Butyl mercaptan. White atoms are hydrogen; black atoms are carbon; and the turquoise atom is selenium. PUBLISHERS RESOURCE GROUP



All three isomers of butyl mercaptan are mobile liquids (liquids that flow easily) with very strong skunk-like odors. At one time, chemists thought that the butyl mercaptans were responsible for the distinctive odor of skunk spray. Research has now shown that some mercaptans are found in skunk spray, but the butyl mercaptans are not among them.

HOW IT IS MADE

A number of methods are available for the production of the butyl mercaptans. The most common method involves the reaction between an iodobutane (a butane molecule with one iodine atom, such as $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{I}$) and a sulfhydryl compound, such as KSH . In this reaction,

Interesting Facts

- The Israeli army reportedly developed a “skunk bomb” containing mercaptans to use as a crowd-control device against Palestinians.
- The Guinness World Records book classifies the butyl mercaptans as being among the two smelliest compounds in the world. The other chemical is ethyl mercaptan.
- The odor of the butyl mercaptans is said to be detectable at a distance of nearly a kilometer (about a half mile) from its source.

the sulphydryl group (-SH) replaces the iodine to form the butyl mercaptan.

COMMON USES AND POTENTIAL HAZARDS

More than 500,000 kilograms (1 million pounds) of the butyl mercaptans are produced each year in the United States. Their primary uses are as solvents and in the preparation of other organic compounds, especially insecticides and herbicides. Among their other uses are:

- Additives for natural gas, which normally has no odor, so that leaks can be easily detected;
- As an additive to enhance the flavor and aroma of certain foods;
- Nutrient in the maintenance of bacterial cultures.

All three isomers of butyl mercaptan are highly toxic. Anyone exposed to their vapors should seek medical attention immediately. The compounds are also very flammable and must be handled with care to prevent their catching fire. Since most individuals never come into contact with the butyl mercaptans, their health and safety risks are of concern primarily to workers who handle them in their everyday jobs.

Words to Know

ISOMER One of two or more forms of a chemical compound with the same molecular formula, but different structural formulas and different chemical and physical properties.

MERCAPTAN Organic compound that contains a sulfhydryl group (-SH) attached to a carbon.

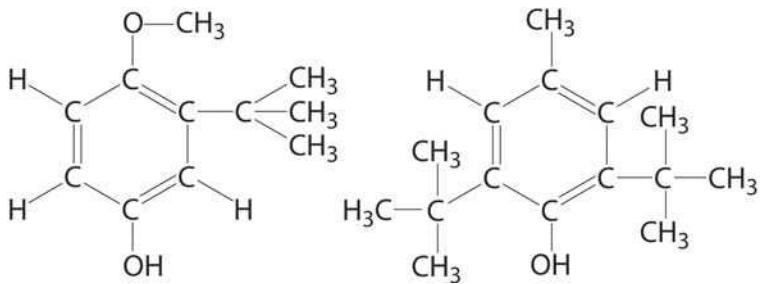
FOR FURTHER INFORMATION

Bauman, Richard. "Getting Skunked: Understanding the Antics behind the Smell." *Backpacker* (May 1993): 30–31.

"n Butyl Mercaptan." International Chemical Safety Cards.
<http://www.cdc.gov/niosh/ipcsneng/neng0018.html> (accessed on September 14, 2005).

"The Stink that Stays." *Popular Mechanics* (December 2004): 26.

Wood, William F. "Chemistry of Skunk Spray." Humboldt State University.
<http://www.humboldt.edu/~wfw2/chemofskunkspray.html> (accessed on September 14, 2005).



OTHER NAMES:
BHA and BHT

FORMULA:
BHA: C₁₁H₁₆O₂; BHT:
C₁₅H₂₄O

ELEMENTS:
Carbon, hydrogen,
oxygen

COMPOUND TYPE:
Organic

STATE:
Solid

MOLECULAR WEIGHT:
BHA: 180.24 g/mol;
BHT: 220.35 g/mol

MELTING POINT:
BHA: 51°C (124°F);
BHT: 71°C (160°F)

BOILING POINT:
BHA: 268°C (514°F);
BHT: 265°C (509°F)

SOLUBILITY:
Both are insoluble in
water and soluble in
ethyl alcohol; BHT is
also soluble in acetone
and benzene

KEY FACTS

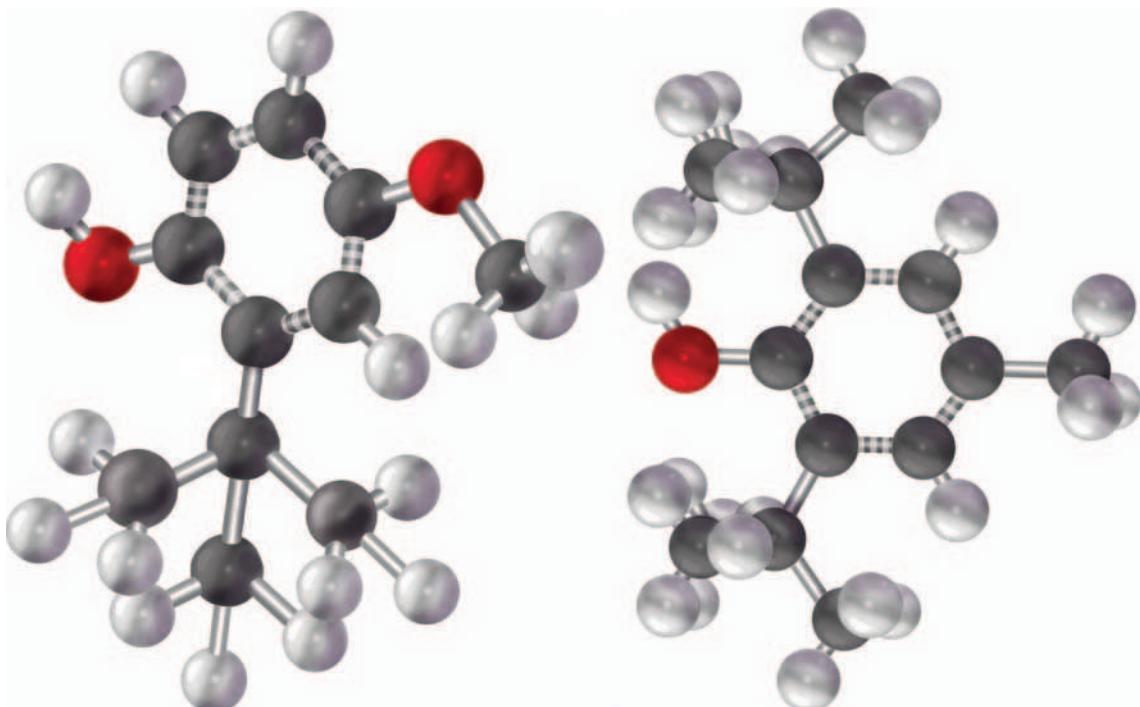
OVERVIEW

Butylated hydroxyanisole (BYOO-til-ay-ted hi-DROK-see-ANN-i-sole) and butylated hydroxytoluene (BYOO-til-ay-ted hi-DROK-see-TOL-yoo-een) are very popular food additives used to preserve fats and oils. They both are antioxidants, which are compounds that prevent oxygen from reacting with substances and changing them into other materials. BHA and BHT prevent the oxidation of fats and oils that would convert them into rancid, foul-smelling, harmful products.

BHA is a white or pale yellow waxy solid with a faint pleasant odor. BHT is a white crystalline solid. Both compounds are members of the phenol family of organic compounds. The phenols are compounds containing a benzene ring of six carbon atoms to which is attached at least one hydroxyl (-OH) group.

HOW IT IS MADE

A variety of methods are available for the preparation of BHA. The most common method involves the reaction



BHA (left) and BHT (right).
Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. Gray sticks indicate double bonds. Striped sticks indicate a benzene ring.

PUBLISHERS RESOURCE GROUP

between *p*-methoxyphenol ($\text{CH}_3\text{CH}_2\text{OC}_6\text{H}_4\text{OH}$) and isobutene ($(\text{CH}_3)_2\text{C}=\text{CH}_2$) or *tert*-butyl alcohol ($(\text{CH}_3)_3\text{COH}$) over a catalyst of silica (silicon dioxide; SiO_2) or alumina (aluminum oxide; Al_2O_3) at temperatures of about 150°C (302°F). BHT is usually made by reacting *p*-cresol ($\text{CH}_3\text{C}_6\text{H}_4\text{OH}$) with isobutene.

COMMON USES AND POTENTIAL HAZARDS

The most important use of both BHA and BHT is as food preservatives. When these compounds are added to products containing fats or oils, oxygen reacts with the additive (BHA or BHT) rather than the food itself, protecting the food from spoiling. Among the vast array of food products containing either BHA or BHT or both are cereals, seasonings, frostings, dessert mixes, instant potatoes, packaged popcorn, baked goods, pie crusts, meat products, potato chips, candy, sausage, freeze-dried meats, butter, cheese, crackers, bread, vegetable oils, margarine, nuts, beer, and chewing gum.

Both compounds are also added to animal feed as preservatives. BHA and BHT are added to a number of non-food products

as well. These products include paraffin wax, lipstick, eye shadow, lip-gloss, mascara, body and face lotions, diaper rash ointment, deodorant soaps, moisturizers, and shaving gels and creams.

The antioxidant properties of BHA and BHT make them suitable for other applications also. For example, they are sometimes added to paints and inks to prevent a "skin" from forming on top of these liquids. The skin is formed when the paint or ink reacts with oxygen in air to form a solid compound. The compounds are also used as additives for the preservation of drugs, rubber products, petroleum products, the plastics used in food wraps, and the petroleum wax coatings used on food boxes.

Questions have been raised about the possible health benefits and risks posed by BHA and BHT. On the one hand, these compounds may contribute to good health by destroying substances in the body that can lead to cancer. They may also destroy the herpes virus and the human immunodeficiency virus (HIV). Most nutritionists agree that antioxidants can be helpful in protecting cells and tissues from aging and damage by oxygen.

On the other hand, BHA and BHT may also be responsible for certain health problems. For example, some people appear to be allergic to these compounds, causing skin rashes, hives, or tightness in the chest, although such reactions are thought to be rare. Other people may have problems metabolizing BHA and BHT properly, resulting in a build-up of the compounds in their bodies.

Scientists are uncertain about the possible long-term health effects of BHA and BHT. Some studies suggest that these compounds may affect liver or kidney function or may be carcinogenic. Studies on this subject are inconclusive, however, and the U.S. Food and Drug Administration has decided that the two compounds can be used as food additives if the amount added is less than 0.02 percent by weight.

As with all chemicals, direct exposure to large doses of BHA or BHT can cause serious health problems. In one experiment, for example, adults who ingested four grams (0.14 ounce) of BHA experienced stomach pain, vomiting, dizziness, confusion, and temporary loss of consciousness. That amount of BHA is tens or hundreds of thousands of times the amount one would consume in foods or other products, however. Only people who work directly with the compound would ever worry about an effect such as this one.

Words to Know

CARCINOGEN A chemical that causes cancer in humans or other animals.

FOR FURTHER INFORMATION

“Butylated Hydroxyanisole.” CHEC’s HealtheHouse.

http://www.checnet.org/healthehouse/chemicals/chemicals_detail2.asp?Main_ID=330 (accessed on January 9, 2006).

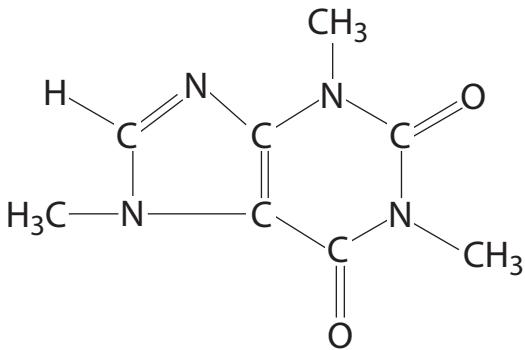
“Butylated Hydroxyanisole.” *Gale Encyclopedia of Science*. Edited by K. Lee Lerner and Brenda Wilmoth Lerner. 3rd ed., vol. 1. Detroit: Gale, 2004.

“Butylated Hydroxyanisole.” Hazardous Substances Data Bank.
<http://toxnet.nlm.nih.gov/> (accessed on Jnauary 9, 2006).

“Butylated Hydroxyanisole (BHA; tert butyl 4 hydroxyanisole) and Butylated Hydroxytoluene (BHT; 2,6 di tert butyl p cresol).” In *Food Antioxidants: Technological, Toxicological, and Health Perspectives*, edited by D. L. Madhavi, S. S. Deshpande, and D. K. Salunkhe. New York: Dekker, 1996.

“Butylated Hydroxytoluene.” *Gale Encyclopedia of Science*. Edited by K. Lee Lerner and Brenda Wilmoth Lerner. 3rd ed., vol. 1. Detroit: Gale, 2004.

“Butylated Hydroxytoluene.” International Chemical Safety Cards.
<http://www.cdc.gov/niosh/ipcsneng/nengo841.html> (accessed on January 9, 2006).



OTHER NAMES:	Methyltheobromine; theine
FORMULA:	$C_8H_{10}N_4O_2$
ELEMENTS:	Carbon, hydrogen, nitrogen, oxygen
COMPOUND TYPE:	Organic base (alkaloid)
STATE:	Solid
MOLECULAR WEIGHT:	194.19 g/mol
MELTING POINT:	238°C (460°F)
BOILING POINT:	Not applicable; begins to sublime at about 90°C (190°F)
SOLUBILITY:	Slightly soluble in water and alcohol; soluble in chloroform

KEY FACTS

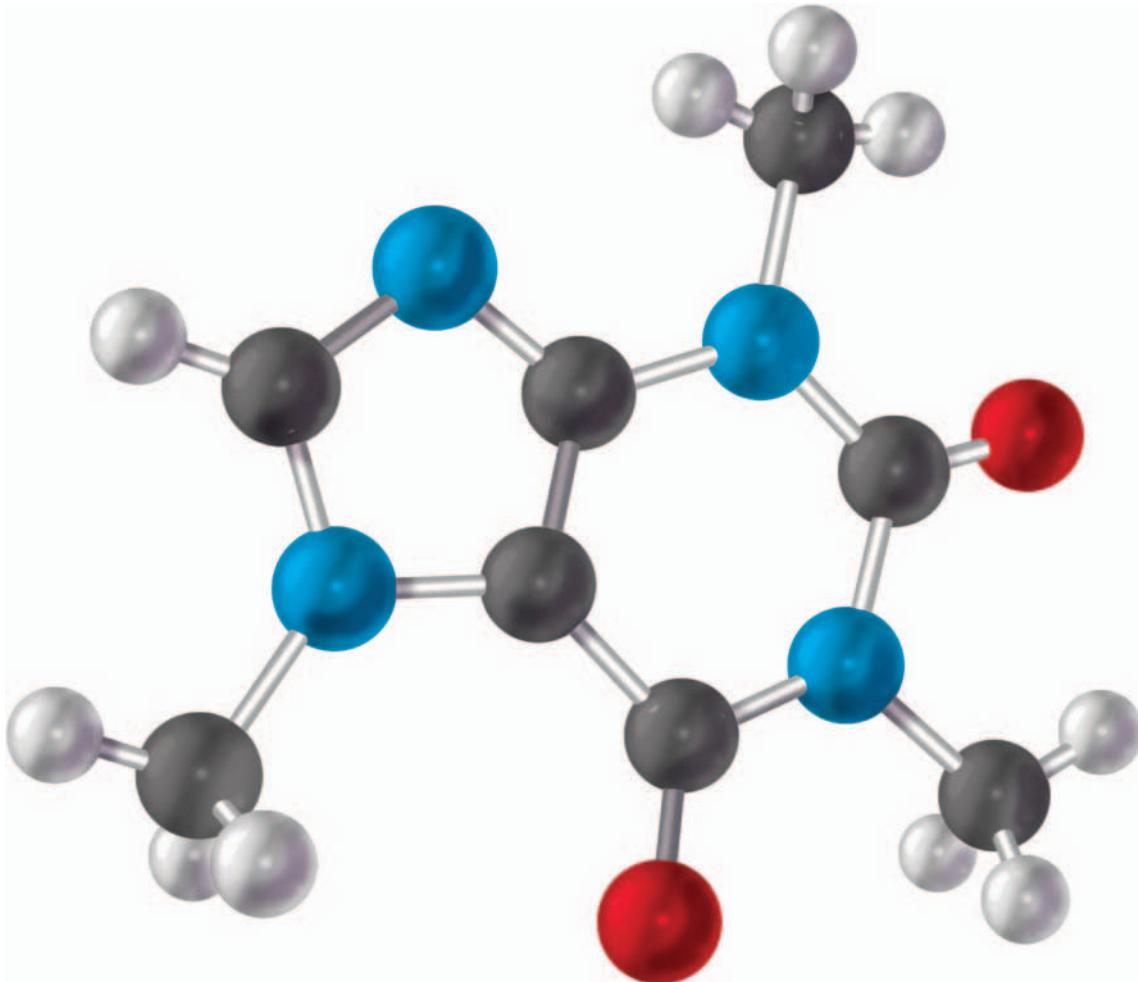
Caffeine

OVERVIEW

Caffeine (kaf-EEN) is an organic base that occurs naturally in a number of plant products, including coffee beans, tea leaves, and kola nuts. It occurs as a fleecy white crystalline material, often in the form of long, silky needles. It usually exists as the monohydrate, $C_8H_{10}N_4O_2 \cdot H_2O$, although it gives up its water of hydration readily when exposed to air.

Scientists believe that humans have been drinking beverages that contain caffeine for thousands of years. The first recorded reference to a caffeine drink can be found in a Chinese reference to the consumption of tea by the emperor Shen Nung in about 2700 BCE. Coffee is apparently a much more recent drink, with the earliest cultivation of the coffee tree dated at about 575 CE in Africa.

Caffeine was first studied scientifically by two French chemists, Joseph Bienaimé Caventou (1795–1877) and Pierre Joseph Pelletier (1788–1842), who were very interested in the chemical properties of the alkaloids. Between 1817



Caffeine. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atoms are nitrogen. Gray sticks are double bonds. PUBLISHERS RESOURCE GROUP

and 1821, Caventou and Pelletier successfully extracted caffeine, quinine, strychnine, brucine, chinchonine, and chlorophyll (not an alkaloid) from a variety of plants. The first synthesis of caffeine was accomplished in 1895 by the German chemist Emil Hermann Fischer (1852-1919), who was awarded the 1902 Nobel Prize in chemistry for his work on the alkaloids.

Caffeine belongs to a class of alkaloids called the methylxanthines. Chocolate, from the cocoa tree *Theobroma cacao* contains another member of the class, theobromine. Both caffeine and theobromine are stimulants, that is, compounds

Interesting Facts

- Americans drank 6.3 billion gallons of coffee, 2.4 billion gallons of tea, and 15.3 billion gallons of
- caffeinated soft drinks in 2003.

that act on the nervous system to produce alertness, excitement, and increased physical and mental activity.

HOW IT IS MADE

Caffeine can be extracted from coffee, tea, and kola plants by one of three methods. These methods are used primarily to produce the decaffeinated counterparts of the products: decaffeinated coffee, decaffeinated tea, or decaffeinated soft drinks. A commercial variation of these procedures is to treat the waste products of tea or coffee processing, such as the dust and sweepings collected from factories, for the extraction of caffeine.

In the first of the three extraction methods, the natural product (coffee beans, tea leaves, or kola beans) are treated with an organic solvent that dissolves the caffeine from the plant material. The solvent is then evaporated leaving behind the pure caffeine. A second method follows essentially the same procedure, except that hot water is used as the solvent for the caffeine. A more recent procedure involves the use of supercritical carbon dioxide for the extraction process. Supercritical carbon dioxide is a form of the familiar gas that exists at high temperature and high pressure. It behaves as both a liquid and a gas. Not only is the supercritical carbon dioxide procedure an efficient method of extracting caffeine, but it has virtually none of the harmful environmental and health problems associated with each of the other two methods of extraction.

Caffeine is also made synthetically by heating a combination of the silver salt of theobromine ($C_7H_8N_4O_2Ag$) with methyl iodide (CH_3I), resulting in the addition of one carbon

Words to Know

ALKALOIDS Organic bases that contain the element nitrogen.

DIURETIC A substance that increases the flow of urine.

WATER OF HYDRATION Water that has combined with a compound by some physical means.

and two hydrogens to the theobromine molecule and converting it to caffeine.

COMMON USES AND POTENTIAL HAZARDS

Caffeine is used in foods and drinks and for medical purposes. Its primary action is to stimulate the central nervous system. People drink coffee, tea, or cola drinks to stay awake and alert because caffeine creates a feeling of added energy. It does this by increasing heart rate, improving blood flow to the muscles, opening airways to aid breathing, and releasing stored energy from the liver to provide fuel for the body. In large quantities, caffeine can also cause nervousness, insomnia, and heart problems. The effects of caffeine can linger in the body for more than six hours. In medical applications, caffeine is sometimes used as a heart stimulant for patients in shock, to treat apnea (loss of breathing) in newborn babies, to counteract depressed breathing levels as a result of drug overdoses, and as a diuretic.

Caffeine stimulates the brain in two ways. First, because it has a chemical structure similar to that of adenosine, it attaches to adenosine receptors in the brain. Adenosine is a substance that normally attaches to those receptors, slowing brain activity and causing drowsiness. By blocking those receptors, caffeine increases electrical activity in the brain, creating a feeling of alertness. Caffeine also works in the brain like drugs such as heroin and cocaine, although in a much milder way. Like those drugs, caffeine increases dopamine levels. Dopamine is a chemical present in the brain that increases the body's feeling of pleasure.

Studies have shown that caffeine can become addictive. People who use the compound eventually need to take more

and more of it to get the same effect. When some people try to stop using caffeine, they may suffer from headache, fatigue, and depression, though these symptoms can be controlled by gradually reducing the amount of caffeine consumed. Either way, withdrawal symptoms end after about a week.

FOR FURTHER INFORMATION

“Caffeine.” Neuroscience for Kids.

<http://faculty.washington.edu/chudler/caff.html> (accessed on September 23, 2004).

“Caffeine Chemistry.” Erowid Vault.

http://www.erowid.org/chemicals/caffeine/caffeine_chemistry.shtml (accessed on September 23, 2005).

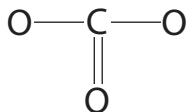
Cherniske, Stephen. *Caffeine Blues: Wake Up to the Hidden Dangers of America's #1 Drug*. New York: Warner Books, 1998.

Kluger, Jeffrey. “The Buzz on Caffeine.” *Time* (December 20, 2004): 52.

Weinberg, Alan Bennet, and Bonnie K. Bealer. *The World of Caffeine: The Science and Culture of the World's Most Popular Drug*. New York: Routledge, 2002.

See Also Carbon Dioxide; Theobromine

Ca



OTHER NAMES:
Limestone; chalk;
aragonite; calcite

FORMULA:
 CaCO_3

ELEMENTS:
Calcium, carbon,
oxygen

COMPOUND TYPE:
Inorganic salt

STATE:
Solid

MOLECULAR WEIGHT:
100.09 g/mol

MELTING POINT:
Calcite: 1330°C
(2430°F); aragonite:
decomposes at about
 825°C (1520°F)

BOILING POINT:
Not applicable

SOLUBILITY:
Very slightly soluble
in water; soluble in
dilute acids; insoluble
in organic solvents

KEY FACTS

OVERVIEW

Calcium carbonate (KAL-see-um CAR-bun-ate) is one of the most common compounds on Earth, making up about 7 percent of Earth's crust. It occurs in a number of minerals and other natural materials, including aragonite, calcite, chalk, limestone, marble, marl, oyster shells, pearls, and travertine. Stalactites and stalagmites found in caves are made primarily of calcium carbonate. As indicated by the melting points of aragonite and calcite, the compound's physical properties may differ somewhat depending on its crystal form. It typically occurs as an odorless, tasteless white powder or colorless crystals.

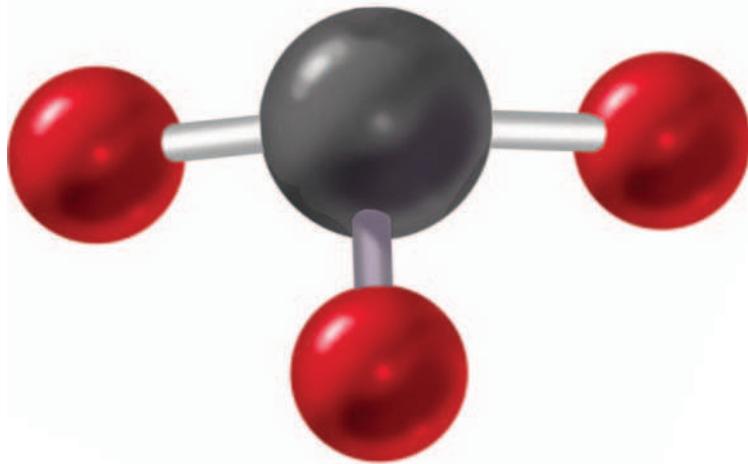
HOW IT IS MADE

Calcium carbonate is so abundant in nature that demand for the compound can be met by mining natural sources, such as chalk, limestone, and marble quarries. The compound can also be produced in the laboratory by reacting



Calcium carbonate. Red atoms are oxygen; black atom is carbon; white atom is calcium.
Gray stick is a double bond.

PUBLISHERS RESOURCE GROUP



calcium chloride (CaCl_2) with sodium carbonate (Na_2CO_3). The calcium carbonate formed in this reaction precipitates out of (separates from) the solution and can be recovered by filtration.

COMMON USES AND POTENTIAL HAZARDS

One well known use of calcium carbonate is as an antacid, a substance that neutralizes excess stomach acid (hydrochloric acid; HCl) and relieves the symptoms of acid indigestion, heartburn, upset stomach, and sour stomach. Large amounts of the chemical are also used for a variety of industrial uses, including:

- In agriculture, where it is used to maintain proper acidity of soil and supply calcium needed by growing plants;

Interesting Facts

- It is said that Cleopatra showed her extravagance by dissolving pearls (which are made of calcium carbonate) in vinegar and drinking the resulting solution.
- Stalagmites and stalactites form in caves when calcium bicarbonate (CaHCO_3) dissolved in groundwater reaches the top of a cave and gives up carbon dioxide (CO_2) to the air. When calcium bicarbonate loses

carbon dioxide, it is converted to calcium carbonate, which precipitates out on the roof of the cave as a stalactite. If the calcium bicarbonate does not lose carbon dioxide until it drips off the roof and falls on the floor of the cave, the calcium carbonate that is formed builds up a stalagmite on the floor of the cave.

- In the paper-making industry, where it is used to make strong products with a very white color, glossy finish, and firm texture for printing and dyeing;
- As a filler in the manufacture of plastics to reduce the cost of production;
- In the construction industry, where it is used in the production of concrete structures, such as paving-stones, tubes, and sewage tanks, in ready-mixed concrete, and in prefabricated elements;
- In the production of paints and other types of coating materials because of its ability to provide weather resistance, protect the coating against corrosion, reduce drying time, and maintain the proper acidity of the coating material;
- In a variety of environmental applications, such as counteracting the increased acidity of lakes and other bodies of water caused by acid precipitation, treating waste gases to remove sulfur and nitrogen oxides that pollute the air; and purifying water and waste water; and
- In the manufacture of lime (calcium oxide; CaO) which, itself, has a host of industrial applications.

Words to Know

PRECIPITATE A solid material that settles out of a solution, often as the result of a chemical reaction.

FOR FURTHER INFORMATION

“Calcium Carbonate.” National Institute for Occupational Safety and Health.

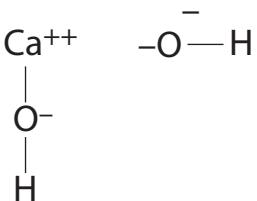
<http://www.cdc.gov/niosh/npg/npgd0090.html> (accessed on September 24, 2005).

“Calcium Carbonate Powder.” Reade.

http://www.reade.com/Products/Minerals_and_Ores/calcium_carbonate.html (accessed on September 24, 2005).

Tegethoff, F. Wolfgang, with Johannes Rohleder and Evelyn Kroker, eds. *Calcium carbonate: From the Cretaceous period into the 21st century*. Boston: Birkhäuser Verlag, 2001.

See Also Calcium Oxide



KEY FACTS	
OTHER NAMES:	Calcium hydrate; caustic lime; hydrated lime; slaked lime
FORMULA:	$\text{Ca}(\text{OH})_2$
ELEMENTS:	Calcium, oxygen, hydrogen
COMPOUND TYPE:	Inorganic base
STATE:	Solid
MOLECULAR WEIGHT:	74.09 g/mol
MELTING POINT:	Not applicable; loses water when heated to become calcium oxide (CaO)
BOILING POINT:	Not applicable
SOLUBILITY:	Slightly soluble in water; soluble in glycerol

Calcium Hydroxide

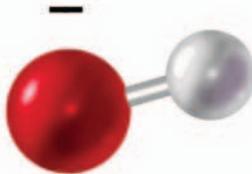
OVERVIEW

Calcium hydroxide (KAL-see-um hye-DROK-side) is a soft, white odorless solid that occurs as granules or a powder. It has a slightly bitter, alkaline taste. Calcium hydroxide absorbs carbon dioxide readily from the air, changing to calcium carbonate (CaCO_3). For this reason, the compound often is contaminated with the carbonate unless it is kept in tightly sealed containers. Calcium hydroxide is a relatively inexpensive chemical and, for that reason, is used in production processes where a base is needed. Although calcium hydroxide is not very soluble in water, a suspension of the finely divided powder in water can be made, a suspension known as limewater or milk of lime.

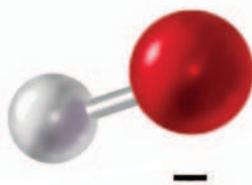
HOW IT IS MADE

Calcium hydroxide is produced by adding water to calcium oxide (CaO). This reaction is highly exothermic and must be carried out with some care to avoid loss of the product through spattering.

Calcium hydroxide. Red atoms are oxygen; white atoms are hydrogen; and the turquoise atom is calcium. PUBLISHERS RESOURCE GROUP



2+



COMMON USES AND POTENTIAL HAZARDS

Calcium hydroxide has a number of uses in building and paving materials. It is found in mortar, plasters, and cements. When used in mortar, calcium hydroxide is mixed with sand and water to make a paste. The paste is put between bricks to hold them together. When water in the paste evaporates, the remaining sand-calcium hydroxide mixture becomes a hard, strong adhesive holding the bricks together.

Calcium hydroxide is also used extensively as a neutralizing agent. As a base, it reacts with acids formed in industrial operations, in the environment, and in other situations. For example, soil that is too acidic for plants to grow properly can be treated with calcium hydroxide. The hydroxide ion (OH^-) in the calcium hydroxide reacts with the hydrogen ion (H^+) in the acidic soil to form water, resulting in a more neutral soil.

Interesting Facts

- One of the oldest methods of coating a surface is with whitewash, a mixture made of calcium hydroxide, water, and chalk (calcium carbonate). Whitewash is used to whiten the walls of a building, a fence, or some other structure.

A similar application involves the treatment of exhaust gases from factories with calcium hydroxide. Nitrogen and sulfur oxides in the gases, which are the primary cause of acid rain, are neutralized when passed through “scrubbers” containing calcium hydroxide.

Among the many other uses of calcium hydroxide are:

- In the leather and hide tanning industry, where calcium hydroxide is used to remove hair from hides and then to neutralize tanning solutions;
- For the purification of sugar solutions, such as those used in making fruit juices;
- As a component of certain types of pesticides;
- As an additive to lubricants and oil drilling fluids, to improve the viscosity of the fluids;
- In the manufacture of paints and waterproof coatings;
- As a food additive, used for the purpose of preventing foods from becoming too acidic;
- As an accelerant (a compound that increases the rate of a chemical reaction) in processes for the production of synthetic rubber and plastics; and
- As a food additive in poultry feed, to improve the strength of egg shells.

Exposure to calcium hydroxide can cause irritation of the skin, eyes, and respiratory system. These conditions are treated satisfactorily by flushing the contaminated area with ample amounts of water. People who work directly with calcium hydroxide in the workplace are more likely to be at risk for health problems because of their exposure to larger

Words to Know

EXOTHERMIC REACTION A reaction in which heat is produced.

amounts of the chemical for longer periods of time. In such cases, more serious health problems, such as burns and blistering of the skin, damage to the retina of the eye, and injury to the gastrointestinal system may develop.

FOR FURTHER INFORMATION

"Calcium Hydroxide or Hydrated Lime." Peters Chemical Company.
http://www.peterschemical.com/calcium_hydroxide_or_hydrated_li.htm (accessed on September 26, 2005).

"Material Safety Data Sheet: Calcium Hydroxide Products." Pulp dent Corporation.
http://www.pulpdent.com/c_hydro/CHMSDS.pdf (accessed on September 26, 2005).

See Also Calcium Oxide

OTHER NAMES:	Lime; quicklime; burnt lime; calx; unslaked lime; fluxing lime
FORMULA:	CaO
ELEMENTS:	Calcium, oxygen
COMPOUND TYPE:	Metallic oxide
STATE:	Solid
MOLECULAR WEIGHT:	56.08 g/mol
MELTING POINT:	2,898°C (5,248°F)
BOILING POINT:	Not available
SOLUBILITY:	Reacts with water to form calcium hydroxide; soluble in acids; insoluble in alcohol and most organic solvents

KEY FACTS

Calcium Oxide

OVERVIEW

Calcium oxide (KAL-see-um OK-side) is an odorless crystalline or powdery solid that, in a pure form, is white to off-gray. It often appears with a yellowish or brownish tint to the presence of impurities, especially iron. Calcium oxide reacts with water to form calcium hydroxide (Ca(OH)_2) with the evolution of significant amounts of heat. The compound is strongly caustic.

Calcium oxide has been known since ancient times. The Roman writer Cato the Elder (234-149 BCE) described one method of producing the compound in 184 BCE. Another early Roman scholar, Pliny the Elder (23-79 CE) discussed the compound at length in his book *Historia Naturalis* (*Natural History*), published in 70 CE. By the early fifteenth century, most of Europe was using calcium oxide (widely referred to as lime) in the construction of buildings. The Scottish chemist Joseph Black (1728-1799) carried out some of the earliest scientific studies of calcium oxide. He found that when the compound is exposed to air, it combines with carbon dioxide to form calcium carbonate.



Calcium oxide. Red atom is oxygen and green atom is calcium. Gray stick shows a double bond. PUBLISHERS RESOURCE GROUP

HOW IT IS MADE

The process for making calcium oxide is believed to be one of the first chemical reactions known to humans, dating back to prehistoric times. When limestone (calcium carbonate; CaCO_3) is heated, carbon dioxide (CO_2) is driven off, leaving calcium oxide behind. The reaction was probably discovered very early in human history because limestone is a common, readily available material in the form of chalk and sea shells, and the amount of heat needed to produce the reaction can easily be produced in a simple wood fire. A more efficient method for carrying out the reaction is to heat the limestone in a kiln (oven) at temperatures of 500°C to 900°C (900°F to $1,600^\circ\text{F}$), resulting in a more complete conversion of calcium carbonate to calcium oxide. This method is still used today for the commercial preparation of calcium oxide.

COMMON USES AND POTENTIAL HAZARDS

The most important single use of calcium oxide is in metallurgy, particularly in the production of steel. When calcium oxide is added to the furnace in which steel is made, it reacts with sulfur, phosphorus, silica, and other impurities present in the mixture from which steel is produced. The complex mixture that results can be poured off the top of the molten steel in the form of a slag, a nonmetallic waste formed during the production of metals. Calcium oxide plays a comparable role in the manufacture of other metals, such as aluminum and magnesium. About 40 percent of all the calcium oxide produced in the United States goes to metallurgical applications.

The next most important use of calcium oxide is in pollution control devices. Smoke that leaves a factory's

Interesting Facts

- Calcium oxide is often used to “lime” lake waters that have been acidified by acid rain. It reacts with and neutralizes acids in the lake formed when nitric and sulfuric acid in acid rain are carried to earth by rain, snow, sleet, and other forms of precipitation.
- When calcium oxide is heated near its melting point, it gives off a brilliant white light. In the years before electricity was available for lighting, particularly during the second half of the nineteenth century, heated lime was used to produce the bright lights needed to illuminate stage productions. From this practice came the expression “being in the limelight” to refer to anyone who was in public view of large groups of people.
- Because it was thought to accelerate the decomposition of soft tissue, quicklime has historically been used in the burial of diseased animals and humans. For example, bodies of plague victims in London in 1666 were directed to be buried in quicklime.

smokestack, for example, contains oxides of sulfur and nitrogen that, in the atmosphere, combine with water to form sulfuric acid and nitric acid. To prevent the formation of these pollutants, “scrubbers” can be installed in smokestacks. The scrubbers contain some chemical that reacts with and neutralizes the oxides of sulfur and nitrogen. One of the most common compounds used for this purpose is calcium oxide. About 15 percent of all calcium oxide used in the United States goes to this application.

Some other uses of calcium oxide include:

- In water treatment plants, to control acidity of the water being treated and to remove impurities present in the water;
- In the construction industry, where it is used to make plaster, mortar, stucco, bricks, and other building materials;
- As a filler to strengthen paper products;

Words to Know

CAUSTIC Strongly basic or alkaline material that irritates or corrodes living tissue.

METALLURGY The science of working with metals.

- As a refractory, a heat-resistant material used to line the insides of furnaces;
- In the production of other chemical materials;
- As an additive for poultry feed;
- In insecticides and fungicides;
- In the removal of hair from hides before tanning; and
- As a food additive to maintain proper acidity and give bulk to a food product.

Exposure to calcium oxide can cause damage to the skin, eyes, nose, and respiratory system. People who use the product in their line of work or at home (for garden purposes, for example) must exercise extreme caution to avoid swallowing, breathing in, or otherwise coming into contact with the chemical. If such contact occurs, it should be washed off completely and medical assistance requested.

FOR FURTHER INFORMATION

“Calcium Oxide or Quicklime.” Peters Chemical Company.

http://www.peterschemical.com/calcium_oxide_or_quicklime.htm (accessed on September 27, 2005).

“Chemical of the Week Lime.”

<http://scifun.chem.wisc.edu/chemweek/Lime/lime.html> (accessed on September 27, 2005).

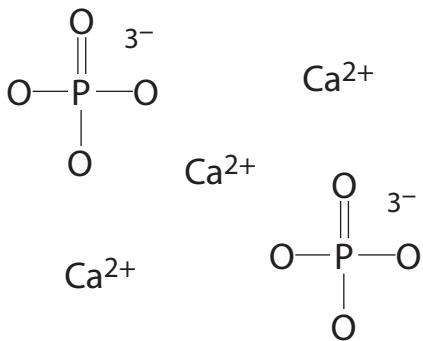
“Cheminfo.” Canadian Centre for Occupational Health and Safety.

<http://intox.org/databank/documents/chemical/calcoxit/cie11.htm> (accessed on September 27, 2005).

“Metallurgy Information Area.” National Lime Association.

<http://www.lime.org/ENVo2/Metal8o2.htm#IS> (accessed on September 27, 2005).

See Also Calcium Carbonate; Calcium Hydroxide



OTHER NAMES:
see Overview

FORMULA:
see Overview

ELEMENTS:
calcium, hydrogen,
phosphorus, oxygen

COMPOUND TYPE:
Inorganic salts

STATE:
Solid

MOLECULAR WEIGHT:
136.06 to 310.20 g/mol

MELTING POINT:
see Overview; tribasic
form: 1670°C
(3040°F)

BOILING POINT:
Not applicable

SOLUBILITY:
See Overview

KEY FACTS

Calcium Phosphate

OVERVIEW

The three forms of calcium phosphate (KAL-see-um FOSS-fate) all occur as tasteless, odorless, colorless to white crystalline or powdery solids.

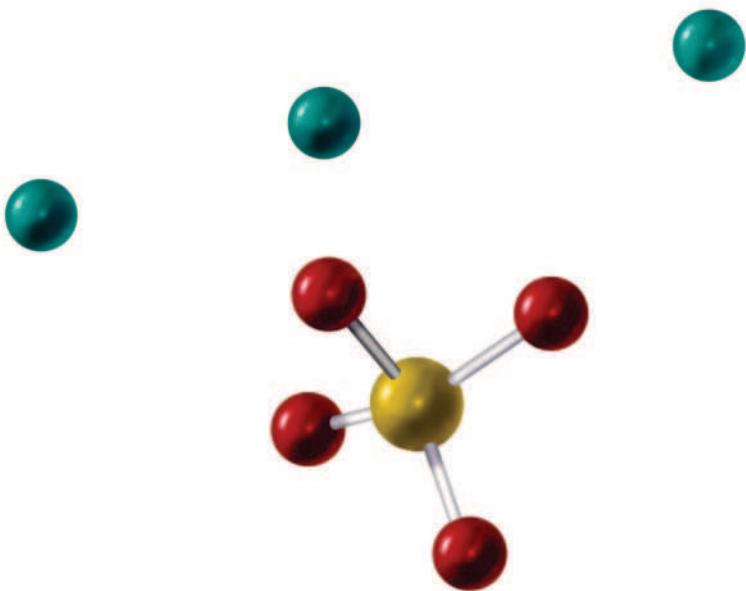
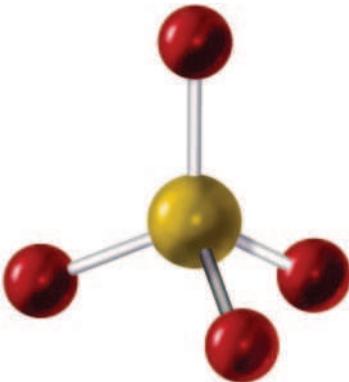
Dibasic calcium phosphate, CaHPO_4 , is also called calcium monohydrogen phosphate, dicalcium orthophosphate, or secondary calcium phosphate. It is usually found in the form of hydrate, such as $\text{CaHPO}_4 \cdot 2\text{H}_2\text{O}$. It does not melt, instead decomposing when heated to 109°C (228°F).

Monobasic calcium phosphate, $\text{Ca}(\text{H}_2\text{PO}_4)_2$, is also known as calcium hypophosphite, calcium biphenylate, acid calcium phosphate, monocalcium orthophosphate, and primary calcium phosphate. It usually exists in the form of the hydrate $\text{Ca}(\text{H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O}$. It decomposes when heated to 200°C (400°F).

Tribasic calcium phosphate, $\text{Ca}_3(\text{PO}_4)_2$, may be called calcium orthophosphate, tricalcium orthophosphate, tertiary calcium orthophosphate, or tricalcium phosphate. Unlike the other two forms, the tribasic form contains no hydrogen atoms.

Calcium phosphate. Red atoms are oxygen; yellow atoms are phosphorus; and the turquoise atoms are calcium. Gray sticks are double bonds.

PUBLISHERS RESOURCE GROUP



The dibasic and tribasic forms of calcium phosphate are insoluble in water and alcohol, but soluble in most acids. Mono-basic calcium phosphate is soluble in water and acids, and insoluble in alcohol. All three compounds are non-flammable.

HOW IT IS MADE

Tribasic calcium phosphate can be obtained directly from rock and minerals such as apatite, a complex and impure form of calcium phosphate, or phosphorite, a mineral that

Interesting Facts

- The forms of calcium phosphate discussed here are only three of many ways in which calcium, phosphorus, and oxygen combine to form calcium phosphate-type compounds. Another such compound is known as hydroxyapatite ($\text{Ca}_5(\text{PO}_4)_3(\text{OH})$), which makes up about 70 percent of an adult tooth.
- Calcium phosphate is one of the two most common compounds found in kidney stones. The second of these compounds is calcium oxalate (CaC_2O_4).

contains calcium phosphate mixed with other compounds. The compound can also be produced synthetically by reacting phosphoric acid (H_3PO_4) with solid calcium hydroxide ($\text{Ca}(\text{OH})_2$), or in the reaction between calcium chloride (CaCl_2) and sodium tripolyphosphate ($\text{Na}_5\text{P}_3\text{O}_{10}$). Dibasic calcium phosphate can also be produced in the former of these synthetic reactions by using an aqueous solution of calcium hydroxide rather than the solid compound. Monobasic calcium phosphate is produced synthetically by adding excess phosphoric acid to a solution of either dibasic or tribasic calcium phosphate and allowing the solution to evaporate.

COMMON USES AND POTENTIAL HAZARDS

Each form of calcium phosphate has its own uses. The dibasic form is used as a nutrient and mineral supplement in animal foods and in certain processed foods, especially cereals. In addition to its nutritional value, dibasic calcium phosphate acts as a dough conditioner, stabilizer, and thickener in foods. The compound is also used in dental products to provide replacement for hydroxyapatite lost to decay or other factors.

The most common application of monocalcium phosphate is in fertilizers, where its function is to provide the phosphorus that plants need to remain healthy and grow properly.

Words to Know

AQUEOUS Consisting of some material dissolved in water.

MORDANT A substance used in dyeing and printing that reacts chemically with both a dye and the material being dyed to help hold the dye permanently to the material.

SYNTHESIS a chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

WATER OF HYDRATION Water that has combined with a compound by some physical means.

It is also used as a food additive in baking powders and wheat flours where it serves to retain the proper acidity of the food product to which it is added. It is also used, like the dibasic form, as a nutritional supplement in processed foods.

Tricalcium phosphate has the largest range of uses of the three forms of the compound, the most important of which are in the manufacture of fertilizers, the production of ceramic materials, and the preparation of other compounds of phosphorus. The compound is also used:

- As a clarifying agent in the purification of sugar solutions;
- As a mordant in the dyeing of cloth;
- In the preparation of dental products;
- As a stabilizer for plastics;
- As a food additive to prevent a powdery product from caking (becoming compacted);
- As a nutritional supplement; and
- As a means of removing radioactive strontium (strontium-90) from milk that has been contaminated by the element.

While none of the forms of calcium phosphate is flammable, they can all irritate the skin, eyes, and respiratory system. Tricalcium phosphate is the most hazardous of the family of compounds. Direct contact with the eyes may cause severe symptoms, including irritation and burning, resulting in damage to the cornea. The compound is also poisonous if taken internally. It can cause stomach pain, vomiting, and low blood pressure. These problems occur when one is exposed

to relatively large amounts of the compound. In the quantities present in foods and other products with which consumers come into contact, the compound poses little risk to users.

FOR FURTHER INFORMATION

“Calcium Phosphate.” Medline Plus.

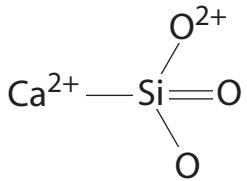
<http://www.nlm.nih.gov/medlineplus/ency/article/oo2889.htm>
(accessed on September 28, 2005).

“Dicalcium Phosphate.” Univar USA

<http://www.msdsvault.org//UNIVARUSA/33222770621C2A2E204EBB8825668Foo61E8ED.OPENDOCUMENT.PDF>
(accessed on September 28, 2005).

“Tricalcium Phosphate.”

<http://www.luminet.net/~wenonah/hydro/3capo.htm> (accessed on September 28, 2005)



OTHER NAMES:
Calcium metasilicate

FORMULA:
 CaSiO_3

ELEMENTS:
Calcium, silicon,
oxygen

COMPOUND TYPE:
Inorganic salt

STATE:
Solid

MOLECULAR WEIGHT:
116.16 g/mol

MELTING POINT:
1540°C (2800°F)

BOILING POINT:
Not applicable

SOLUBILITY:
Insoluble in water and
most other common
solvents

K E Y F A C T S

OVERVIEW

The term calcium silicate (KAL-see-um SILL-i-kate) applies to a number of related compounds consisting of calcium, silicon, and oxygen, all of which tend to occur in a hydrated form. The most common of these forms is calcium metasilicate, for which data are given above. Calcium silicates with the formulas Ca_2SiO_4 and Ca_3SiO_5 are also well known.

Calcium silicate is a component of many minerals, including afwillite, akermanite, andradite, calcite, centrallasite, crestmoreite, diopside, eaklite, grammite, gyrolite, hillebrandite, larnite, and wollastonite. In its pure form, calcium metasilicate is a white to off-white color capable of absorbing up to two and a half times its weight of water. In this form, the hydrated powder retains its ability to flow freely. Addition of a mineral acid, such as hydrochloric acid (HCl) results in the formation of a gel.

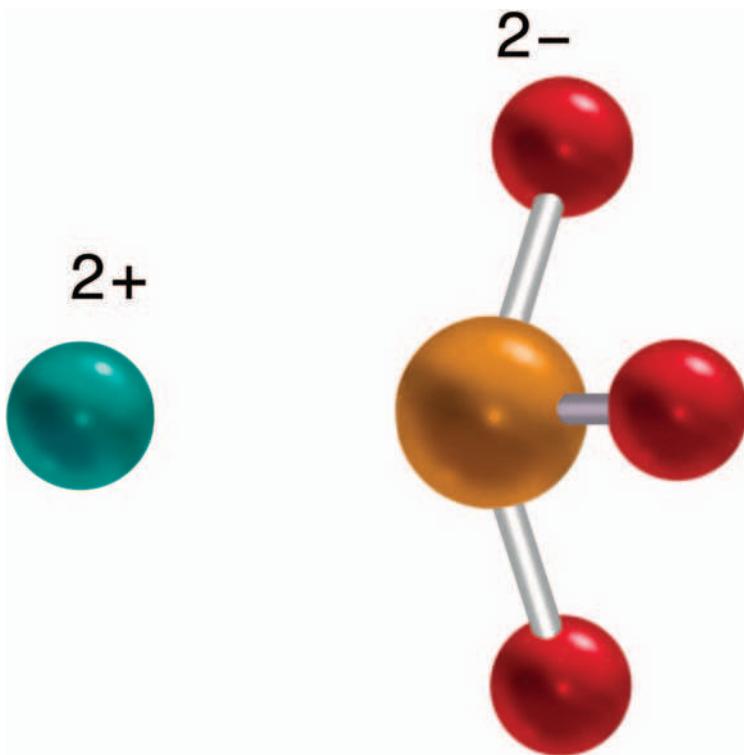
HOW IT IS MADE

Given its abundance in rocks and minerals, calcium metasilicate and its cousins can be extracted from a variety of

Calcium silicate. Red atoms are oxygen; turquoise atom is calcium; orange atom is silicon.

Gray stick is a double bond.

PUBLISHERS RESOURCE GROUP



natural sources. Nonetheless, the compound is usually prepared synthetically in order to ensure that it has the purity required for most commercial applications. The usual method of preparation involves the addition of lime (calcium oxide; CaO) to diatomaceous earth, a material consisting of the fossil remains of single-celled algae known as diatoms. The lime provides the calcium and the diatomaceous earth provides the silicon dioxide (SiO_2) required to make calcium silicate.

COMMON USES AND POTENTIAL HAZARDS

The calcium silicates have a number of uses in industry. Among the most important applications are their uses in building materials, such as some types of glass and cement (especially Portland cement), bricks and tiles for roofs, fire-proof ceilings, and building boards. The compound is also

Interesting Facts

- One of the most common building materials in use today is Portland cement, made from a mixture of oxides, including lime (calcium oxide), sand (silicon dioxide), and aluminum, iron, and magnesium oxides. When these components are mixed and water is added, a complex

set of chemical reactions occur in which calcium silicates are among the major products formed. These calcium silicates are responsible for the great strength for which Portland cement products are known.

used as a filler in the manufacture of paper and some types of plastics, where it gives body to the final product.

Other applications of the calcium silicates include the following:

- As an absorbent for liquids and gases in industrial processes;
- As a food additive, where it absorbs moisture and prevents a product from “caking” (becoming compacted);
- In road construction;
- As a refractory material in industrial furnaces;
- For the insulation of pipes and conduits;
- In the manufacture of paints and other types of coating materials;
- In the preparation of certain types of cosmetics;
- In some fertilizers and insecticides; and
- In some pharmaceutical products, such as antacids.

There have been no reports of calcium silicate's being hazardous or harmful to humans or experimental animals. As with all chemicals, overexposure to large amounts of the compound could result in irritation of the skin, eyes, and respiratory system.

Words to Know

HYDRATE Chemical compound formed when one or more molecules of water is added physically to the molecule of some other substance.

REFRACTORY Having a high melting point, resistant to melting. Refractory materials

are often used to line the interior of industrial furnaces.

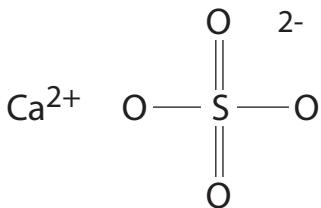
SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

FOR FURTHER INFORMATION

“Occupational Safety and Health Guideline for Calcium Silicate, Synthetic.” Occupational Health and Safety Administration.
<http://www.osha.gov/SLTC/healthguidelines/calciumsilicate/recognition.html#healthhazard> (accessed on September 29, 2005).

“Scientific Principles.” Materials Science and Technology Teacher’s Workshop.
<http://matse1.mse.uiuc.edu/~tw/concrete/prin.html> (accessed on September 29, 2005).

“Wollastonite.” CeramicMaterials.Info.
<http://ceramic materials.com/cermat/material/1705.html> (accessed on September 29, 2005).



OTHER NAMES:

Anhydrous gypsum;
also see Overview for
synonyms of
hydrates

FORMULA:

CaSO_4

ELEMENTS:

Calcium, sulfur,
oxygen

COMPOUND TYPE:
Salt (inorganic)

STATE:
Solid

MOLECULAR WEIGHT:
136.14 g/mol

MELTING POINT:
1460°C (2660°F)

BOILING POINT:
Not applicable

SOLUBILITY:
Insoluble in water and
most organic
solvents

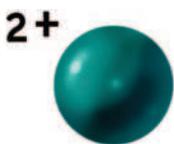
K
E
Y
F
A
C
T
S

OVERVIEW

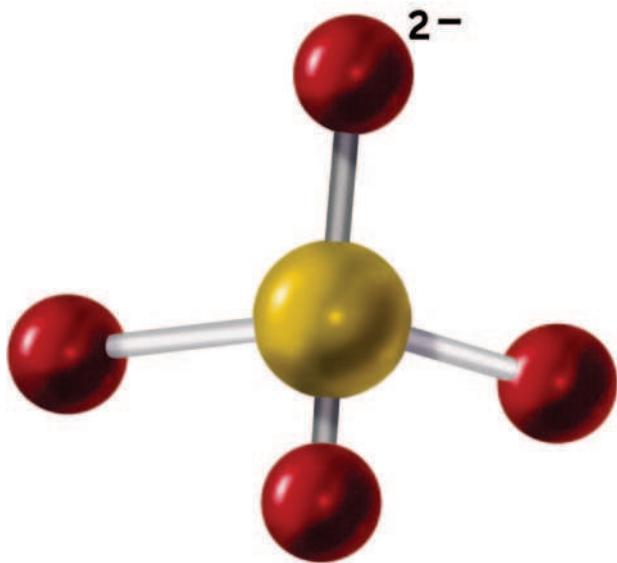
Calcium sulfate (KAL-see-um SUL-fate) occurs in three forms:

- anhydrous calcium sulfate (CaSO_4); the anhydrous form of calcium sulfate is available in two forms, known as insoluble anhydrite and soluble anhydrite;
- calcium sulfate dihydrate ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$), also known as mineral white, terra alba, light spar, precipitated calcium sulfate, native calcium sulfate, and by other names;
- calcium sulfate hemihydrate ($\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$), also known as plaster of Paris and dried gypsum;

The physical properties of the three forms of calcium sulfate differ somewhat from each other, but their chemical properties are essentially the same. Anhydrous calcium sulfate and calcium hemihydrate are fine white odorless powders or crystalline solids, while the dihydrate may occur either as a powder or as white lumps. Both hydrates are



Calcium sulfate. Red atoms are oxygen; turquoise atom is calcium; and yellow atom is sulfur. Gray sticks show double bonds. PUBLISHERS RESOURCE GROUP



converted to the anhydrous form upon heating as, for example: $\text{CaSO}_4 \cdot 2\text{H}_2\text{O} \rightarrow \text{CaSO}_4 + 2\text{H}_2\text{O}$.

Anhydrous calcium sulfate is essentially insoluble in water. As their names suggest, the soluble form of the compound (soluble anhydrite) is somewhat more soluble in water than is the insoluble form (insoluble anhydrite). The dihydrate and hemihydrate are only slightly soluble in water. When water is added to the hemihydrate, a reaction occurs that results in the formation of a hard, solid mass (plaster of Paris) used in making casts, such as those used to hold broken bones in place. Neither the anhydrous form of calcium sulfate or the dihydrate reacts with water in this way.

Both the anhydrous and dihydrate forms of calcium sulfate occur naturally in the form of the minerals anhydrite, angelite, muriacite, and karstenite (CaSO_4); and gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$). These minerals have been known to humans and used by them for thousands of years. The method for converting natural gypsum to the hemihydrate (plaster of

Paris) has also been known and used for a very long period of time. Archaeologists have learned that the Egyptians developed a method for converting gypsum to plaster of Paris, which was then used as mortars to join blocks in buildings, more than 5,000 years ago.

HOW IT IS MADE

Gypsum is an abundant mineral providing a ready natural supply of calcium sulfate. The mineral usually consists of a mixture of the anhydrous and dihydrate forms of calcium sulfate, which can be separated into its component parts. Clay, sand, limestone, and other impurities are also present in most gypsum deposits. The process of separation and purification begins by crushing natural gypsum and heating it in an open kettle to temperatures of 100°C to 125°C (212°F to 257°F) for a few hours. Heating converts the gypsum to the hemihydrate or anhydrous calcium sulfate depending on the temperature of the reaction. The solubility of the final product can also be controlled by the temperature used in the heating container and the time during which the material is heated.

Large amounts of calcium sulfate are also produced as the by-product of other reactions or by synthetic processes. The production of phosphoric acid (H_3PO_4) from sulfuric acid (H_2SO_4) and phosphate rock, for example, results in the formation of calcium sulfate, which can be recovered and purified. Calcium sulfate can also be produced in the reaction between other calcium compounds, such as calcium hydroxide ($Ca(OH)_2$) and sulfuric acid.

COMMON USES AND POTENTIAL HAZARDS

The most widely used form of calcium sulfate is the dihydrate, gypsum, which is an important raw material in the construction industry. It is used in the manufacture of Portland cement, in specialized plasters (known as gypsum plasters) for walls, in the production of wallboard, and in cement blocks and mortars. Gypsum is also used extensively in agriculture as a *conditioning agent* that adds both calcium ions (Ca^{2+}) and sulfate ions (SO_4^{2-}) to the soil. The compound is also used as a raw material in the synthesis of other calcium compounds and in the production of plaster of Paris.

Interesting Facts

- Calcium sulfate has been used in China for more than 2,000 years to thicken soy milk in the production of tofu.
- One reason that calcium sulfate hemihydrate is called plaster of Paris is that large deposits of gypsum (from which plaster of Paris is made) exist near the city of Paris, France. These deposits were long mined to obtain the hemihydrate form of calcium sulfate.
- Ancient Romans used plaster of Paris to cast copies of Greek statues.

The anhydrous form of calcium sulfate also has a number of practical applications, the most important of which are in the manufacture of cement and as a filler in the production of paper. A filler adds body to paper, making it firmer, brighter, and easier to write, draw, and print on. Soluble anhydrite is used as a desiccant, which is a material that removes water from other substances. It is usually sold under the trade name of Drierite®.

The primary use of calcium hemihydrate is, of course, in the production of plaster of Paris. Plaster of Paris has a number of important applications, such as:

- In the construction of arts and crafts projects, such as masks, ceramics, and pottery;
- For the production of plaster casts used to immobilize broken bones;
- As a building material in the manufacture of stucco, drywall, division panels, ceiling tablets, plaster board, and wall plaster.

Some other important applications of calcium sulfate include:

- As a firming agent in foods such as canned vegetables, soft-serve ice cream, regular ice cream, frostings, and gelatins;
- As an additive to animal feeds to provide the calcium and sulfate that animals need in their diets;

Words to Know

DESICCANT A material that removes water from other substances.

from simple beginning chemicals, or reactants.

SYNTHESIS A chemical reaction in which some desired chemical product is made

- In the production of polishing powders;
- As a paint pigment (white); and
- In a variety of metallurgical processes, such as the conversion of zinc minerals to zinc metal.

Although nontoxic, calcium sulfate can irritate the respiratory tract if inhaled. It may cause symptoms such as coughing, shortness of breath, and nosebleeds. The compound can also irritate the skin and eyes, causing redness and pain. Ingestion of the compound can cause stomach pains, nausea, and vomiting.

FOR FURTHER INFORMATION

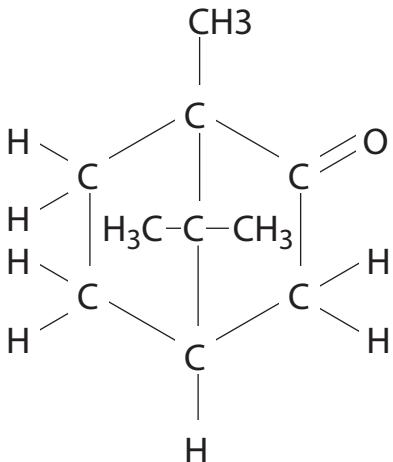
“A Brief History of Plaster and Gypsum.” Association of Lifecasters International.
http://www.artmolds.com/ali/history_plaster.html (accessed on December 10, 2005).

“Calcium Sulfate.” National Organic Program. Agricultural Marketing Service.
<http://www.ams.usda.gov/nop/NationalList/TAPReviews/CaSO4.pdf> (accessed on December 10, 2005).

“Calcium Sulfate, Anhydrous, Powder.” J. T. Baker.
<http://www.jtbaker.com/msds/englishhtml/co497.htm> (accessed on December 10, 2005).

Dontsova, Katerina, et al. “Gypsum for Agricultural Use in Ohio Sources and Quality of Available Products.” Ohio State University Extension Fact Sheet.
<http://ohioline.osu.edu/anr/fact/0020.html> (accessed on December 10, 2005).

“Gypsum Statistics and Information.” U.S. Geological Survey.
<http://minerals.usgs.gov/minerals/pubs/commodity/gypsum/> (accessed on December 10, 2005).



OTHER NAMES:
Gum camphor; also see Overview.

FORMULA:
 $C_{10}H_{16}O$

ELEMENTS:
Carbon, hydrogen, oxygen

COMPOUND TYPE:
Cyclic ketone (organic)

STATE:
Solid

MOLECULAR WEIGHT:
152.23 g/mol

MELTING POINT:
178.3°C to 178.8°C (352.9°F to 353.4°F); varies depending on isomer

BOILING POINT:
207.4°C (405.3°F; depends on isomer); some forms sublime before boiling

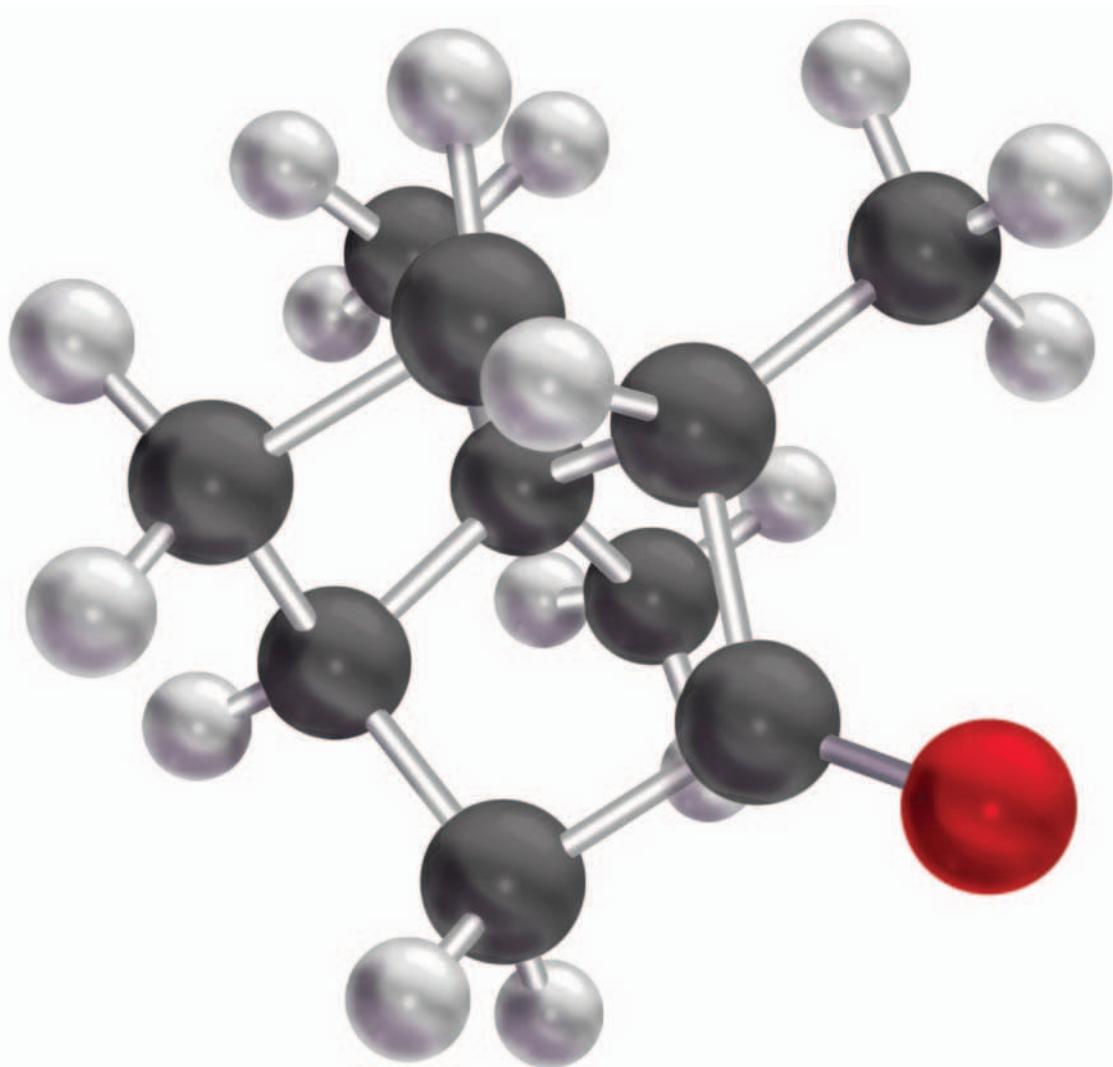
SOLUBILITY:
Insoluble in water; soluble in alcohol, ether, acetone, benzene, and other organic solvents

KEY FACTS

OVERVIEW

Camphor (KAM-for) is also known as gum camphor; 2-camphanone; 1,7,7-trimethylbicyclo[2.2.1]heptan-2-one; and 1,7,7-trimethylbicyclo[2.2.1]-2-heptanone, among others. It is a volatile white waxy substance with a strong, characteristic odor and a bitter, cooling taste. Its odor has been described as fragrant, aromatic, pungent, or penetrating, similar to that of mothballs. The compound is obtained from the camphor tree, *Cinnamomum camphora*, which is native to many parts of the world, including the Indonesian islands of Java and Sumatra, as well as China, Taiwan, Japan, and Brazil.

The first mention of camphor by a Westerner occurs in the writings of the Venetian explorer Marco Polo (1254-1324) in the late thirteenth century. The Chinese had apparently been using camphor for many centuries in the embalming of their dead and the manufacture of soap, among other applications. The Japanese also had a long history of using the compound in the preparation of varnish and ink, to dilute paint, to burn for light, and to protect clothing from moths.



Camphor. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon. PUBLISHERS RESOURCE GROUP

Camphor was also used throughout Asia as a medicine to treat a wide variety of ailments, from upset stomach to rheumatism.

HOW IT IS MADE

Camphor can be obtained from every part of the *Cinnamomum camphora* tree, including the trunk, branches, roots, and leaves. The compound is extracted from tree parts by

Interesting Facts

Despite its name, the camphor tree, *Cinnamomum camphora*, does not produce cinnamon. That compound

comes from a related member of the family, *Cinnamomum zeylanicum*.

steam distillation and crystallization of the camphor. The camphor thus obtained is one of two possible isomers, the dextrorotatory (or "+") form of the compound. (The other isomer is the levorotatory or "-" form.) Since the natural sources of camphor are inadequate to meet commercial demand, camphor is also made synthetically. The process begins by converting the straight-chain hydrocarbons in turpentine oil ($C_{10}H_{16}$) to a cyclic isomer, α-pinene ($\alpha\text{-C}_{10}H_{16}$), which is then oxidized to make camphor. Synthetic camphor differs from natural camphor in that it consists of both isomers of the compound, the dextrorotatory (+) and levorotatory (-) forms. This difference has no effect on the chemical properties of camphor, although it has some modest effect on its biological properties.

COMMON USES AND POTENTIAL HAZARDS

Camphor has a long history of medicinal use. People have used it as a liniment to treat muscle aches and arthritis; as a topical anesthetic; as a treatment for nervousness and hysteria; to treat diarrhea, bronchitis, and infectious fevers; and even as an injection to cure heart failure. Very little medical evidence exists for most of these treatments. As a result, its primary medicinal use today is in lotions and ointments for the relief of mild pain and discomfort due to arthritis, rheumatism, and sore muscles. Some common products that contain camphor are Absorbine Arthritic Pain Lotion®, Ben-Gay®, Campho-phenique®, Panalgesic®, Sloan's Liniment®, and Vicks VapoRub®. In these preparations, the compound acts as both a mild analgesic (pain-reliever) and antiseptic (an agent to prevent the growth of disease-causing microorganisms). Its pronounced odor may also help open nasal passages. Camphor is also an important ingredient in

Words to Know

ANALGESIC A substance that relieves pain.

DISTILLATION A process of separating two or more substances by boiling the mixture of which they are composed and condensing the vapors produced at different temperatures.

ISOMERS Two or more forms of a chemical compound with the same molecular formula, but different structural formulas and different chemical and physical properties.

some antipruritic agents, substances used to relieve the discomfort of itching.

Although camphor is probably best known to most people because of its medicinal uses, the compound also has a number of important industrial applications. For example, it is added to cellulose nitrate as a plasticizer, a substance added to plastic to make it softer and more flexible. Other uses for camphor include:

- As a moth and mildew repellent in mothballs and moth flakes;
- In the embalming of dead bodies;
- In the manufacture of fireworks and explosives;
- As a preservative in pharmaceuticals and cosmetics;
- In the production of lacquers and varnishes; and
- As a chemical intermediary in the production of a number of organic compounds.

One of camphor's most serious disadvantages as a medicinal product is that it is poisonous. Drinking straight camphorated oil can kill a person. In smaller doses, it can cause mental confusion, irritability, jerky movements of the arms and legs, neuromuscular hyperactivity, and seizures. Symptoms can appear between five and ninety minutes after swallowing the camphor. In 1980, the United States Food and Drug Administration limited the amount of camphor allowed in consumer products to 11 percent. It also banned products that contain concentrated camphor, such as camphorated oil and camphorated liniment.

FOR FURTHER INFORMATION

“Camphor.”

<http://www.innvista.com/health/herbs/camphor.htm> (accessed on September 29, 2005).

“Camphor.” Botanical.com.

<http://www.botanical.com/botanical/mgmh/c/campho13.html> (accessed on September 29, 2005).

“Camphor; Gum Camphor; Laurel Camphor. *Camphora Officinalis*.” Drugstore Museum.

http://www.drugstoremuseum.com/sections/level_info2.php?level_id=71&level=2 (accessed on September 29, 2005).

See Also Cellulose Nitrate

OTHER NAMES:
Carbonic anhydride; carbonic acid gas
FORMULA:
CO_2
ELEMENTS:
Carbon, oxygen
COMPOUND TYPE:
Nonmetallic oxide
STATE:
Gas
MOLECULAR WEIGHT:
44.01 g/mol
MELTING POINT:
Not applicable; liquefies under pressure at $-56.56^{\circ}C$ ($-69.81^{\circ}F$)
BOILING POINT:
Sublimes at $-78.4^{\circ}C$ ($-109^{\circ}F$)
SOLUBILITY:
Soluble in water; slightly soluble in alcohol and some other organic solvents

KEY FACTS

Carbon Dioxide

OVERVIEW

Carbon dioxide (KAR-bun dye-OK-side) is a colorless, odorless, tasteless, non-combustible gas that can also exist under pressure as a clear, colorless, odorless, tasteless liquid and as a white, snow-like solid commonly known as dry ice. When dry ice is warmed it sublimes (passes directly from the solid to the gaseous state without first melting) at $-78.4^{\circ}C$ ($-109^{\circ}F$).

The true nature of carbon dioxide was discovered over an extended period of time beginning with the research of the Flemish physician and chemist Jan Baptista van Helmont (1580-1635?). In about 1603, van Helmont isolated a gas produced during the combustion of wood and proved that it was distinct from air. At the time, air was generally regarded as an element that could not be divided into separate components. Van Helmont called the gas *gas sylvestre* ("wood gas"), a substance we now know to be carbon dioxide. Credit for understanding the true nature of carbon dioxide also goes to the Scottish chemist Joseph Black (1728-1799) who produced



Carbon dioxide. Red atoms are oxygen and black atom is carbon. Gray sticks show double bonds. PUBLISHERS RESOURCE GROUP

carbon dioxide by heating calcium carbonate (CaCO_3). Black called the gas fixed air and conducted the first extensive studies of its properties.

The first practical use for carbon dioxide was discovered in the mid-eighteenth century by the English chemist Joseph Priestley (1733–1804). Priestley found that passing carbon dioxide into water produced a sparkling, refreshing drink that he predicted would one day become a great commercial success. He was, of course, correct, since water containing carbon dioxide is the basic component of which all soda drinks are made.

HOW IT IS MADE

Carbon dioxide is produced in nature by a number of reactions. Among the most common is the combustion (burning) of the fossil fuels (coal, oil, and natural gas). The gas is also produced during the decay of organic material, the fermentation of carbohydrates by yeast, and the respiration of animals. In the laboratory, the simplest and most direct method of preparation is to treat a carbonate, such as calcium carbonate, with an acid, such as hydrochloric acid (HCl).

Carbon dioxide is obtained commercially as the by-product of a number of industrial reactions. For example, when calcium carbonate is heated to produce lime (CaO), carbon dioxide is released and captured as a by-product. The steam reforming (refining) of petroleum results in the production of a mixture of gases known as synthesis gas, consisting of carbon dioxide, carbon monoxide, hydrogen, and nitrogen. Carbon dioxide can be separated from the other components of synthesis gas for commercial uses. Carbon dioxide also produces as a by-product of the manufacture of ammonia (NH_3) by the Haber-Bosch process.

Interesting Facts

- Carbon dioxide is the fourth most abundant gas in the atmosphere (after nitrogen, oxygen, and argon) with a concentration of about 0.036 percent. Researchers have found that the concentration of carbon dioxide in the atmosphere has been increasing at a regular rate for at least the last forty years. They believe that the reason for this increase is the escalating use of fossil fuels by humans to heat homes and offices; drive cars, trucks, trains, and airplanes; and to power industrial operations. They further suspect that an increase in the amount of carbon dioxide in the atmosphere may have significant long-term effects on the planet's climate.
- Scientists have become very interested in a form of carbon dioxide known as supercritical carbon dioxide (SCCO₂ or SC-CO₂). Under the proper conditions of temperature and pressure, carbon dioxide (as SCCO₂) behaves as both a liquid and a gas at the same time. This property has proved to be very valuable in using SCCO₂ as a highly efficient solvent that has no environmental disadvantages.

COMMON USES AND POTENTIAL HAZARDS

Carbon dioxide plays an essential role in most biological processes that take place on Earth's surface. Plants use carbon dioxide as a raw material to make the carbohydrates on which their structures are based. When animals eat plants, those carbohydrates are then used to build and maintain their body structures.

In addition to its role in natural processes, carbon dioxide has many commercial and industrial applications. One of the most important uses is in the carbonation of beverages. Although beers and sparkling wines contain carbon dioxide from natural sources (the fermentation of sugars by yeasts), nearly all carbonated beverages have their carbon dioxide added artificially. The carbon dioxide adds a zesty taste to the beverage and helps to preserve it.

Words to Know

SUBLIME Changing of states from solid to gas without becoming liquid first.

SYNTHESIS GAS A mixture of several gases, such as carbon dioxide and hydrogen, used

to produce compounds such as methanol and ammonia that can be separated out from the synthesis gas.

Carbon dioxide is also used as a fire extinguishing agent. Its use for this purpose is based on the facts that it does not burn itself and is heavier than air. Thus, when sprayed on a fire, carbon dioxide settles down on top of the flames and prevents oxygen from reaching the burning material. The carbon dioxide can be supplied in a variety of ways in a fire extinguisher. In some devices, carbon dioxide gas is produced as the result of a chemical reaction that occurs within the fire extinguisher. In other devices, liquid carbon dioxide is released from the extinguisher.

Carbon dioxide is also used in gaseous, liquid, or solid form as a refrigerant. As a gas, it is used as the “working fluid” in refrigerators, the fluid that circulates through the refrigerator changing back and forth from gas to liquid, absorbing heat in the process. In the form of dry ice, carbon dioxide is a very efficient and convenient method for cooling objects to very low temperatures (close to the sublimation point of carbon dioxide, about -78.4°C (-109°F)).

Some other uses of carbon dioxide include the following:

- As an aerosol propellant;
- To provide an oxygen-free atmosphere in which to conduct welding and other operations with flammable materials;
- In the industrial manufacture of carbonates;
- For cloud seeding to promote modifications in the weather (increases or decreases in rain fall);
- In the fumigation of rice to preserve the product for extended periods of time;
- As an artificial smoke in theater productions;

- As a moderator to slow down the speed of neutrons traveling in a nuclear power plant;
- In the frozen food industry;
- To enrich the air in a greenhouse, providing additional carbon dioxide to promote plant growth; and
- For the hardening of foundry molds and cores.

In general, carbon dioxide poses little or not threat to humans in concentrations to which one is normally exposed. Dry ice may pose a hazard if not handled carefully as its very low temperature can cause damage to the skin.

FOR FURTHER INFORMATION

“Carbon Dioxide (CO₂) Applications and Uses.” Universal Industrial Gases, Inc.

<http://www.uigi.com/carbon dioxide.html> (accessed on September 29, 2005).

“Carbon Dioxide, Oxygen, and the Air.” Skool.ie.

http://www.skool.ie/skool/examcentre_jc.asp?id=1980 (accessed on September 29, 2005).

“Chemical of the Week: Carbon Dioxide.”

[http://scifun.chem.wisc.edu/chemweek/CO₂/CO₂.html](http://scifun.chem.wisc.edu/chemweek/CO2/CO2.html) (accessed on September 29, 2005).

See Also Calcium Carbonate; Calcium Oxide

OTHER NAMES:

None

FORMULA:

CO

ELEMENTS:

Carbon, oxygen

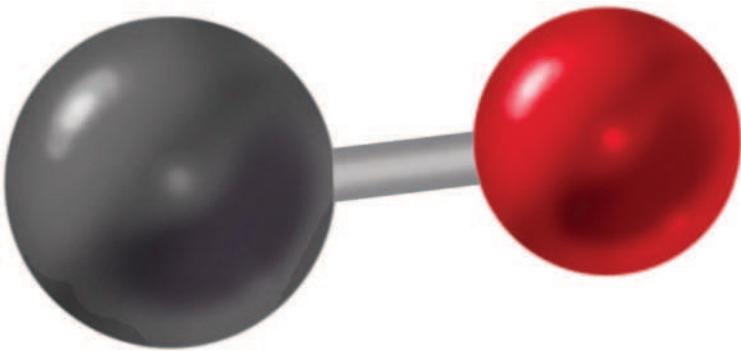
COMPOUND TYPE:
Nonmetallic oxide**STATE:**

Gas

MOLECULAR WEIGHT:
28.01 g/mol**MELTING POINT:**−205.02°C
(−337.04°F)**BOILING POINT:**
−191.5°C (−312.7°F)**SOLUBILITY:**
Slightly soluble in water; soluble in alcohol and chloroform**K E Y F A C T S****OVERVIEW**

Carbon monoxide is a colorless, odorless, tasteless, toxic gas. It is one of the most common poisons in the environment and is responsible for thousands of deaths and hospital emergency room visits each year in the United States. Carbon monoxide is produced from fuel-burning appliances, such as space heaters, furnaces, stoves, and vehicles. It is also a component of cigarette smoke. Carbon monoxide is flammable and capable of forming an explosive mixture with air.

Under ideal circumstances, a carbon-containing fuel such as charcoal, natural gas, or wood burns in air to form carbon dioxide and water. Lacking an adequate supply of oxygen, at low temperatures, or under certain other conditions, incomplete combustion occurs. Incomplete combustion is a process in which a fuel is not completely oxidized. In such instances, carbon monoxide is produced along with carbon dioxide and water. An inadequate supply of oxygen can occur in enclosed, poorly vented spaces, such as the interior of houses, garages, and cars.



Carbon monoxide. Red atom is oxygen and black atom is carbon. Stick is a double bond.

PUBLISHERS RESOURCE GROUP

The discovery of carbon monoxide is usually credited to the English chemist Joseph Priestley (1733–1804). Between 1772 and 1799, Priestley investigated the properties of carbon monoxide and recognized the difference between carbon monoxide and carbon dioxide. The gas was first prepared synthetically by the French chemist Joseph Marie François de Lassone (1717–1788) in 1776, although he mistakenly identified it as hydrogen. The correct chemical formula for carbon monoxide was first identified by the English chemist William Cruikshank (1745–1800) in 1800.

HOW IT IS MADE

A number of methods are available for the commercial production of carbon monoxide. In one procedure, air is passed over hot coke, graphite, or anthracite coal to make producer gas, a mixture of carbon dioxide, carbon monoxide, hydrogen, and water vapor. In a similar procedure, steam is passed over hot coke or graphite to make water gas, a mixture of carbon monoxide, carbon dioxide, hydrogen, and nitrogen. In a third procedure, steam is mixed with natural gas to form synthesis gas, consisting of carbon dioxide, carbon monoxide, and hydrogen. In all three procedures, the carbon monoxide component of the gas mixture produced in the reaction can be separated out from the other gases.

Yet a fourth method for making carbon monoxide involves the partial oxidation of hydrocarbon gases obtained from natural gas or petroleum. These gases consist of carbon-hydrogen compounds which, when oxidized, are converted to carbon monoxide, carbon dioxide, and water. By controlling

Interesting Facts

- Some authorities believe that the American writer Edgar Allan Poe (1809–1849) owed his wild imagination and early death to chronic carbon monoxide poisoning, caused by the gas lighting used in his home.
- Public health authorities estimate that the number of suicides resulting from the inhalation of carbon monoxide fumes from automobiles has decreased by about 80 percent since the introduction of catalytic converters in cars. Catalytic converters reduce the amount of carbon monoxide released in a car's exhaust.
- Carbon monoxide bonds to hemoglobin in the blood stream 200 times as efficiently as does oxygen. This ability to exclude oxygen from the blood is responsible for the toxic effects caused by carbon monoxide in the body.

the amount of reactants used and the conditions of the reaction (temperature and pressure), the portion of carbon monoxide produced can be increased.

COMMON USES AND POTENTIAL HAZARDS

Carbon monoxide has three major industrial uses. The first is in the synthesis of a large variety of organic compounds. For example, it takes part in a group of reactions known as the Fischer-Tropsch reactions in which carbon monoxide is first reduced with hydrogen gas and then converted to any number of organic compounds that contain oxygen. The gas is also used to make acetic acid, a major industrial chemical used in the synthesis of polymers and other organic products.

A second application of carbon monoxide, either by itself or in conjunction with other gases, is as an industrial fuel. The gas burns very efficiently with the release of large amounts of heat and relatively few undesirable by-products (the most important being carbon dioxide).

Words to Know

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

REDUCTION Chemical reaction in which oxygen is removed from a substance or electrons are added to a substance.

SYNTHESIS A chemical reaction in which some desired chemical product is made

from simple beginning chemicals, or reactants.

SYNTHESIS GAS A mixture of several gases, such as carbon dioxide and hydrogen, used to produce compounds such as methanol and ammonia that can be separated out from the synthesis gas.

A third industrial use for carbon monoxide is in the refining of metals. Most metal ores exist in the form of oxides or sulfides when extracted from the earth. For example, the two most important ores of iron are magnetite (Fe_3O_4) and hematite (Fe_2O_3). After an ore has been mined, it is treated to remove the oxygen or sulfur in the ore to obtain a pure metal. Carbon monoxide is often used for this purpose with oxide ores because it combines with oxygen from the ore to form carbon dioxide, leaving the metal behind: $\text{Metal oxide} + \text{CO} \rightarrow \text{Metal} + \text{CO}_2$.

Carbon monoxide is a highly toxic gas that has noticeable health effects even in relatively small concentrations. When the concentration of carbon monoxide reaches levels of about 100 ppm (parts per million), an individual is likely to experience mild headaches, fatigue, shortness of breath, and errors in judgment. As the concentration of carbon monoxide increases, these symptoms become more pronounced. Exposure to a concentration of more than 400 ppm for more than three hours is likely to put a person at serious health risk. He or she may begin to lose consciousness and experience serious disorientation. At concentrations of more than 1,500 ppm, death is likely in less than an hour. These symptoms vary somewhat depending on a person's age and overall health.

Most governmental agencies have set a recommended limit of 35 ppm for periods of up to eight hours. For purposes

of comparison, the normal concentration of carbon monoxide in the atmosphere in an open area tends to be less than 1 ppm. But in urban areas or other locations with many vehicles, gas heaters, wood burning stoves, or other sources of carbon monoxide, carbon monoxide levels can be much higher. Someone traveling inside a car on a busy freeway, for example, may be exposed to carbon monoxide levels of up to 25 ppm. Concentrations of up to 100 ppm have been measured in the center of busy downtown urban areas.

FOR FURTHER INFORMATION

"Carbon Monoxide." Chemical Fact Sheet. Plastics and Chemicals Industries Association.

http://www.pacia.org.au/_uploaditems/docs/3.carbon_monoxide.pdf (accessed on October 2, 2005).

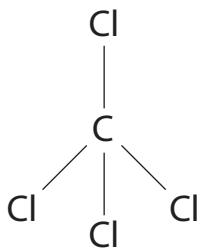
"Carbon Monoxide Questions and Answers." Consumer Product Safety Commission. CPSC Document #466.

<http://www.cpsc.gov/cpscpub/pubs/466.html> (accessed on October 3, 2005).

Dwyer, Bob, et al. *Carbon Monoxide: A Clear and Present Danger*. Mount Prospect, Ill.: ESCO Press, 2004.

Patnaik, Pradyot. *Handbook of Inorganic Chemicals*. New York: McGraw Hill, 2003, 1887 191.

See Also Acetic Acid; Carbon Dioxide



KEY FACTS	
OTHER NAMES:	Tetrachloromethane; perchloromethane
FORMULA:	CCl_4
ELEMENTS:	Carbon, chlorine
COMPOUND TYPE:	Halogenated hydrocarbon; alkyl chloride (organic)
STATE:	Liquid
MOLECULAR WEIGHT:	153.82 g/mol
MELTING POINT:	-22.62°C (-8.72°F)
BOILING POINT:	76.8°C (170°F)
SOLUBILITY:	Insoluble in water; miscible with ethyl alcohol, ether, benzene, chloroform, and most other organic solvents

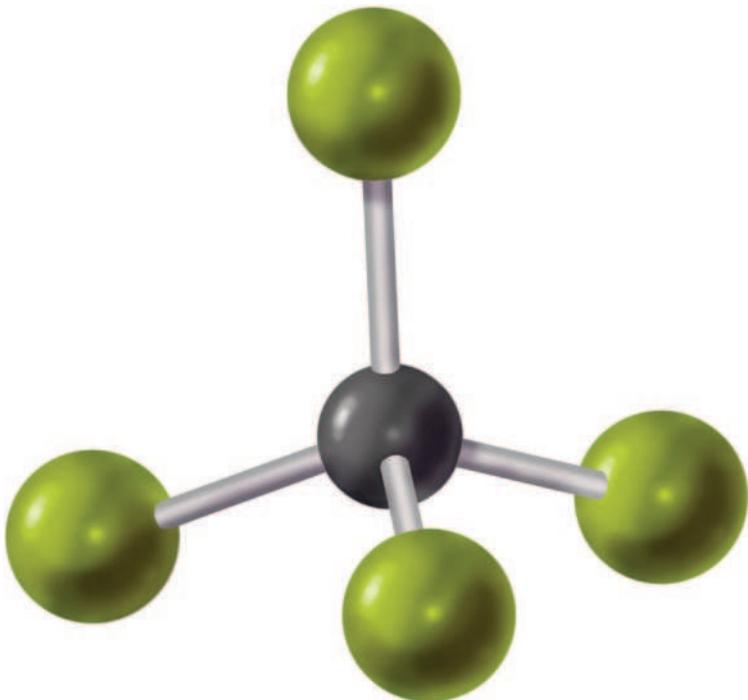
Carbon Tetrachloride

OVERVIEW

Carbon tetrachloride (CAR-bun tet-ra-KLOR-ide) is a colorless, nonflammable liquid with a sweetish odor and a density 1.5 times that of water. The compound is classified as a halogenated hydrocarbon because all of the hydrogens in methane (CH_4), a hydrocarbon, have been replaced by halogen atoms (chlorine, in this case). Carbon tetrachloride was first prepared in 1839 by German-born French chemist Henri Victor Regnault (1810-1878). Regnault made carbon tetrachloride by reacting chloroform (trichloromethane; CHCl_3) with chlorine: $\text{CHCl}_3 + \text{Cl}_2 \rightarrow \text{CCl}_4 + \text{HCl}$.

Until the mid-twentieth century, carbon tetrachloride had a number of important applications, including use as a dry cleaning fluid, a refrigerant, and an industrial and chemical solvent. It was also used in fire extinguishers. By the 1950s, however, evidence of the serious health hazards posed by carbon tetrachloride began to accumulate, and its use was dramatically cut. In 1970, the U.S. Food and Drug Administration banned its use in consumer products in the United States.

Carbon Tetrachloride. Black atom is carbon; green atoms are chlorine. All sticks are single bonds. PUBLISHERS RESOURCE GROUP



Production of carbon tetrachloride was not seriously affected in the long run by this ban, however, as a new use was found for the chemical: the manufacture of chlorofluorocarbons (CFCs). Between the 1930s and the 1970s, the chlorofluorocarbons were among the most successful and widely used of all synthetic organic compounds. They were used in a number of applications, including as refrigerants, blowing agents, solvents, and cleaning agents. They were also used in the production of a variety of organic compounds. In the 1950s and the 1960s, demand for carbon tetrachloride was increasing at a rate of about 10 percent a year in response to the demand for CFCs.

Eventually, that use of carbon tetrachloride was also lessened. Studies conducted by American chemists Mario Molina (1943-) and F. Sherwood Rowland (1927-) showed that CFCs were responsible for the destruction of the Earth's ozone layer, posing potentially serious health hazards to plants and animals living on the planet. In response to this discovery, the world's nations agreed to cut back on the

Interesting Facts

- Before the ban imposed by the Montreal Protocol, more than 90 percent of the chloroform produced was used in the manufacture of chlorofluorocarbons.
- Although carbon tetrachloride was once used as a fire extinguisher, this use had its potential risks. When used on electrical fires, carbon tetrachloride can react with oxygen in the air to produce phosgene (CCl_2O), a poisonous gas used by Germany armies during World War I (1914–1918).
- The carcinogenic effects of carbon tetrachloride are now thought to result from the fact that the compound is converted by enzymes in the body to phosgene, which then attack cells and genes, initiating the processes by which cancer develops.
- Researchers have discovered that the bacterium *Shewanella oneidensis* is capable of degrading carbon tetrachloride, suggesting a possible simple and safe way to clean up areas contaminated with the compound.

production of CFCs and the compounds from which they are made. According to the Montreal Protocol on Substances That Deplete the Ozone Layer, adopted in 1987, the production of carbon tetrachloride for use as a raw material in the manufacture of CFCs was to be first reduced, and then banned entirely by the year 2000. As a result, the production of carbon tetrachloride and its use in modern society has diminished considerably.

HOW IT IS MADE

Carbon tetrachloride is now made by a method invented more than a century ago. Carbon disulfide (CS_2) is treated with chlorine gas, producing sulfur monochloride (S_2Cl_2) and carbon tetrachloride: $\text{CS}_2 + 3\text{Cl}_2 \rightarrow \text{CCl}_4 + \text{S}_2\text{Cl}_2$.

This process is very efficient because the sulfur monochloride can then be treated with additional carbon

CARBON TETRACHLORIDE

disulfide to produce more carbon tetrachloride: $\text{CS}_2 + 2\text{S}_2\text{Cl}_2 \rightarrow \text{CCl}_4 + 6\text{S}$.

A relatively small amount of carbon tetrachloride is also made in the reaction between methane gas (CH_4) and chlorine gas at temperatures of 250°C to 400°C (500°F to 750°F): $\text{CH}_4 + 4\text{Cl}_2 \rightarrow \text{CCl}_4 + 4\text{HCl}$.

COMMON USES AND POTENTIAL HAZARDS

The effect of bans on the production of carbon tetrachloride in recent decades has been dramatic. Production of the compound in the United States dropped from more than 50,000 metric tons (55,000 short tons) in 1989 to less than 2 metric (2.2 short tons) in 1998. The compound cannot be used in any consumer or household products and has only a limited number of industrial applications. In general, its use is restricted to those situations in which no satisfactory substitute has been found. Some of these uses include:

- Certain specific and limited types of refrigeration systems;
- Degreasing metals;
- The manufacture of certain electronic products;
- Additives for some kinds of gasoline;
- The recovery of tin from tin plating operations;
- The fumigation of limited types of grains;
- Solvents for a number of industrial operations.

The average person is likely to come into contact with carbon tetrachloride in one of three ways: by inhaling vapors of the compound, by ingesting (swallowing) the compound or any product in which it is an ingredient, and (to a lesser extent), through the skin. The primary health risks from exposure to carbon tetrachloride involve damage to the liver and kidneys and, at high doses, the central nervous system. Some symptoms associated with exposure to carbon tetrachloride include soreness of the liver or kidneys, nausea, headaches, drowsiness, and disorientation. Although strong evidence is lacking, the U.S. Environmental Protection Agency has labeled carbon tetrachloride as a probable carcinogen.

Words to Know

CARCINOGEN A chemical that causes cancer in humans or other animals.

MISCIBLE Able to be mixed; especially applies to the mixing of one liquid with another.

OZONE LAYER A region of Earth's atmosphere that contains a large quantity of ozone, which absorbs some of the sun's radiation.

FOR FURTHER INFORMATION

"Carbon Tetrachloride." 11th Report on Carcinogens. National Toxicology Program.

<http://ntp.niehs.nih.gov/ntp/roc/eleventh/profiles/s029carb.pdf> (accessed on December 29, 2005).

"Carbon Tetrachloride." International Programme on Chemical Safety.

<http://www.inchem.org/documents/ehc/ehc/ehc208.htm#SubSectionNumber:3.2.1> (accessed on December 29, 2005).

"Carbon Tetrachloride." Matheson Tri Gas.

<http://www.mathesontrigas.com/msds/MATo4310.pdf> (accessed on December 29, 2005).

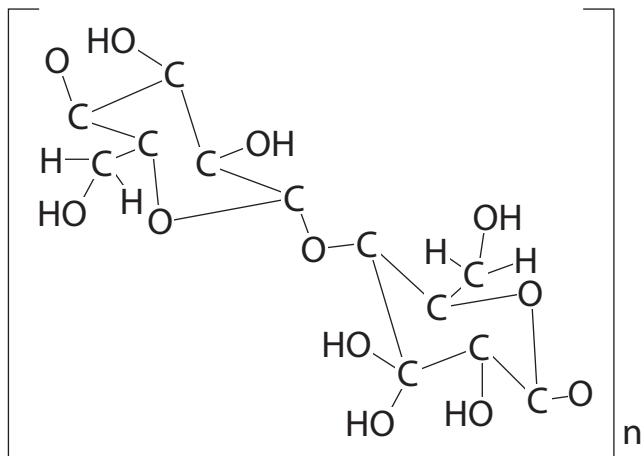
"Carbon Tetrachloride." National Safety Council.

<http://www.nsc.org/library/chemical/carbone.htm> (accessed on December 29, 2005).

"Chemical Fact Sheet: Carbon Tetrachloride." Spectrum Laboratories.

<http://www.speclab.com/compound/c56235.htm> (accessed on December 29, 2005).

See Also Chloroform



OTHER NAMES:	None
FORMULA:	$(C_6H_{10}O_5)_n$
ELEMENTS:	Carbon, hydrogen, oxygen
COMPOUND TYPE:	Polysaccharide (carbohydrate polymer; organic)
STATE:	Solid
MOLECULAR WEIGHT:	Very large, in excess of 100,000 g/mol
MELTING POINT:	Decomposes at temperatures above 260°C (500°F)
BOILING POINT:	Not applicable
SOLUBILITY:	Insoluble in water and all common organic solvents

KEY FACTS

OVERVIEW

Cellulose (SELL-you-lohs) is a colorless to white, tasteless, odorless, polysaccharide fiber found in the cell walls of all land plants and some bacteria, seaweeds, algae, and fungi. Polysaccharides (the term means “many sugars”) are polymers consisting of monosaccharide (simple sugar) monomers joined together in very large molecules. The monomer of which cellulose is made is glucose, also known as blood sugar, dextrose, or grape sugar. The subscript “n” at the end of the chemical formula indicates that a large number of these monomers combine to make the polymer. Cellulose provides the structural support for plants and other organisms in which it occurs. It is generally regarded as the most common organic compound found in nature.

Cellulose never occurs in nature in a pure form. Cotton is the purest natural form, consisting of about 90 percent cellulose. Flax (linen fiber) consists of about 70 to 75 percent cellulose, wood of about 40 to 50 percent cellulose, and seaweeds and algae of about 25 to 30 percent cellulose.



Cellulose. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. PUBLISHERS RESOURCE GROUP

Cellulose was first discovered in 1819 by the French naturalist Henri Braconnot (1781-1855). The compound was first isolated and analyzed fifteen years later by the French botanist Anselme Payen (1795-1871), who gave it its modern name of cellulose based on its origin ("cell") plus the suffix -ose. The earliest chemical studies of cellulose were conducted by a team of English chemists, Charles Frederick Cross (1855-1935), Edward John Bevan (1856-1921), and Clayton Beadle (1868-1917), who identified the compound we now know as cellulose and reported on its structure and properties in the 1890s and early 1900s.

Cellulose, in its natural forms such as wood, cotton, and linen, has countless numbers of uses as food, clothing, construction materials, and other applications. It also serves as

Interesting Facts

- Scientists estimate that plants worldwide synthesize up to one trillion metric tons of cellulose annually.

the raw material in making derivatives such as cellulose acetate, cellulose nitrate, and cellulose xanthate with many other uses and applications.

HOW IT IS MADE

Cellulose is synthesized in plants and some microorganisms through the process known as photosynthesis. In that process, carbon dioxide (CO_2) and water (H_2O) are combined in a complex series of reactions to produce glucose ($\text{C}_6\text{H}_{10}\text{O}_5$) and oxygen (O_2). Glucose molecules are then linked to each other to form successively larger and more complex molecules, eventually resulting in the formation of cellulose.

Commercially, most cellulose is extracted from wood by one of two methods, the kraft (sulfate) process or the steam explosion process. The product of these reactions is wood pulp, which consists primarily of cellulose. In the kraft process, wood chips are treated with a solution of sodium hydroxide (NaOH) and sodium sulfide (Na_2S) at temperatures of about 175°C (350°F) for two to six hours. This process usually results in a yield of about 40 to 45 percent wood pulp. The pulp is then treated with a bleaching agent, such as calcium or sodium hypochlorite ($\text{Ca}(\text{OCl})_2$ or NaClO) or chlorine dioxide (ClO_2) to remove the color of lignin and other impurities.

In the steam explosion process, wood chips are saturated with moisture and then exposed to temperatures of 200°C to 250°C (400°F to 500°F) at pressures of one to five atmospheres. In this process, cellulose fibers are physically separated from lignin fibers, which are the other major constituent of wood. In either process, yields of up to 99 percent pure cellulose can be obtained.

Words to Know

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

COMMON USES AND POTENTIAL HAZARDS

The number of products made from pure cellulose is almost endless. The largest volume of such products include paper and paper products, including (with 2001 production numbers): newsprint (5.8 million metric tons; 6.4 million short tons), printing and writing paper (22.1 million metric tons; 24.4 million short tons), packaging and related uses (3.9 million metric tons; 4.3 million short tons), tissue paper (6.4 million metric tons; 7.0 million short tons), containerboard (26.6 million metric tons; 29.3 million short tons), and boxboard (7.7 million metric tons; 8.5 million short tons). Other uses of cellulose include:

- Manufacture of cotton products, such as items of clothing, sheeting, and industrial fabrics;
- Production of other organic products, especially ethanol (ethyl alcohol; grain alcohol) and methanol (methyl alcohol; wood alcohol);
- As insulation and soundproofing;
- As a food additive, where it is used to thicken and add bulk to food products;
- In equipment used in analytical chemistry, such as chromatographic devices used to separate the components of a mixture.

A major industrial use of cellulose is in the preparation of various cellulose derivatives, primarily cellulose acetate, cellulose nitrate, and cellulose xanthate, each of which has a number of applications.

FOR FURTHER INFORMATION

“Cellulose.” Fibersource.

http://www.fibersource.com/f_tutor/cellulose.htm (accessed on September 30, 2005).

“Cellulose.” London South Bank University.

<http://www.lsbu.ac.uk/water/hycel.html> (accessed on September 30, 2005).

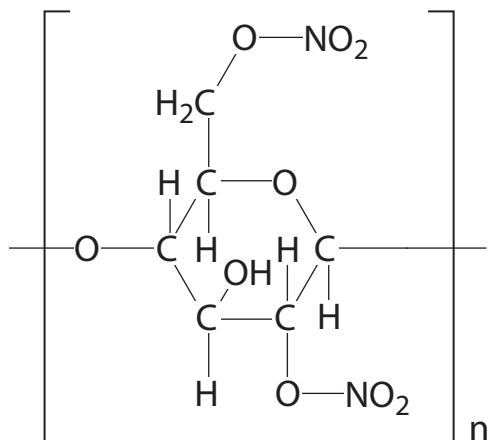
“Cellulose Processing.” Organic Materials Review Institute.

http://www.omri.org/cellulose_final.pdf (accessed on September 30, 2005).

Cross, Charles, Clayton Beadle, Edward Bevan. *Cellulose. An Outline of the Chemistry of the Structural Elements of Plants, with Reference to Their Natural History and Industrial Uses.* Elibron. Replica of 1895 edition by Longmans, Green, and Co., London.

http://www.elibron.com/english/other/item_detail.phtml?msg_id=10006054 (accessed on September 30, 2005).

See Also Cellulose Acetate; Cellulose Nitrate; Cellulose Xanthate



OTHER NAMES:
Nitrocellulose;
nitrocotton;
guncotton; pyroxylin

FORMULA:
 $\text{C}_{12}\text{H}_{16}\text{N}_4\text{O}_{18}$

ELEMENTS:
Carbon, hydrogen,
nitrogen, oxygen

COMPOUND TYPE:
Carbohydrate poly-
mer (organic)

STATE:
Solid

MOLECULAR WEIGHT:
Varies, but very large

MELTING POINT:
Not applicable;
ignites at about
 170°C (340°F)

BOILING POINT:
Not applicable

SOLUBILITY:
Insoluble in water,
alcohol, ether, and
most organic
solvents; also see
Overview

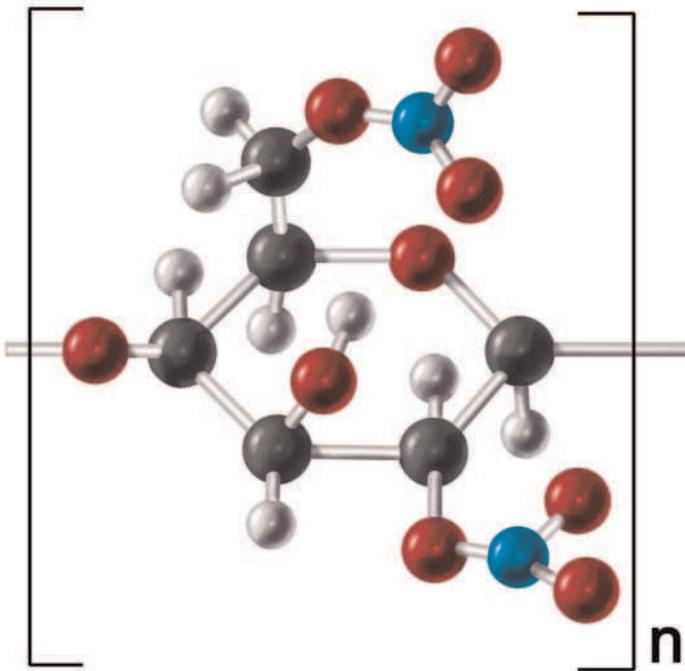
KEY FACTS

OVERVIEW

Cellulose nitrate (SELL-you-lohs NYE-trate) is a derivative of cellulose made by adding nitric acid (HNO_3) to cellulose. Its physical properties tend to differ somewhat depending on the relative amount of nitrogen present in the compound. For example, the so-called “high nitrogen” form is soluble in both acetone but insoluble in an alcohol/ether mixture. The “low-nitrogen” form is soluble in both solvents. Cellulose nitrate is available in a variety of forms, ranging from colorless to white flakes or powder to a liquid or semi-solid gel-like material.

Cellulose nitrate was first discovered in 1845 by the Swiss-German chemist Christian Friedrich Schönbein (1799–1868). The story is told that Schönbein used his wife’s cotton apron to clean up some nitric acid that he had spilled on his laboratory floor. He was amazed to discover that the cotton and nitric acid reacted to form a new compound that exploded when heated, releasing a puff of black smoke. He immediately recognized the potential application for the new

Cellulose nitrate. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atoms are nitrogen. PUBLISHERS RESOURCE GROUP



material as an explosive and suggested the name guncotton for it. He developed a more efficient method for making the product and, within a short period of time, guncotton had become very popular as an explosive for both construction and military purposes.

Researchers soon found that modifying the way in which cellulose nitrate is prepared could significantly change its chemical and physical properties, and, therefore, its uses. For example, the British inventor Alexander Parkes (1813–1890) experimented with using a variety of solutions to dissolve cellulose nitrate, after which the material could be precipitated in one or another form. At first he called the reformulated cellulose nitrate “synthetic ivory,” because of its similarity to natural ivory, but later referred to the product as parkesine, named after himself. He was unable to obtain financing needed for the commercial manufacture of his invention, however, and it fell to an American inventor, John Wesley Hyatt (1837–1920) to reinvent a similar product less than a decade later. Hyatt and his brother Isaiah developed a mixture of cellulose nitrate and camphor that was an even more

Interesting Facts

- Cellulose nitrate was first used for the manufacture of photographic film in 1889 by the Eastman Kodak company. Films made in that era are highly unstable, with a tendency to catch fire spontaneously. Efforts are now under way to protect and preserve some classic motion pictures originally captured on cellulose nitrate film that are unlikely to survive for many more years without such treatment.
- Although the Hyatt billiard balls made out of celluloid were very popular at first, they had one serious disadvantage: They had a tendency to explode upon contact with each other.
- Items of clothing made from cellulose nitrate in the late nineteenth century were sometimes known as “mother-in-law” clothing. They were so flammable that some people thought that the item of clothing along with a cigarette lighter was the perfect gift for an unappreciated mother-in-law.

satisfactory substitute for ivory than was Parke's parkesine. The Hyatts obtained a patent for their new product, to which they gave the name celluloid. Originally developed to meet a growing demand for an ivory-like material that could be used to make billiard balls, celluloid eventually had a number of other applications, including false teeth and photographic film.

HOW IT IS MADE

Cellulose nitrate is produced in essentially the same way as it was (albeit accidentally) by Christian Schönbein in 1845. A mixture of nitric acid and sulfuric acid are added to cellulose in one of various forms (as wood pulp, powder, flakes, or a colloidal mixture with water, for example). The properties of the product formed depend on the relative quantities of each reagent used and the conditions under which the reaction is carried out. One of the most important variables is the amount

of nitrogen present in the cellulose nitrate formed. The more nitrogen present, the less stable the compound is and the more appropriate it is for use as an explosive. The less nitrogen present, the more stable the compound, and the more suitable it is for household and commercial applications.

COMMON USES AND POTENTIAL HAZARDS

Cellulose nitrate continues to be used as an explosive, most commonly in a product known as smokeless gunpowder. The name comes from the fact that, when ignited, the compound produces less smoke than is the case with most other explosives. Smokeless gunpowder is used primarily for smaller weapons, such as artillery rockets and projectiles. A solution of cellulose nitrate dissolved in ether, known as collodion, has been used for a number of applications since its discovery independently in 1846 by the French chemist Louis-Philippe Ménard (dates not available) and the American physician Dr. J. P. Maynard (dates not available). These uses include the coating of wounds; preparation of bandages, surgical dressings, and other medical paraphernalia; the production of photographic film; and the manufacture of laboratory equipment and devices. Cellulose nitrate is also used as a coating material in fast-drying automobile lacquers, on the cloth used to cover hardbound books, and as a leather covering.

Cellulose nitrate poses two major health risks. First, old film made of cellulose nitrate has a tendency to deteriorate and decompose. That process results in the release of oxides of nitrogen, which may be toxic to a greater or lesser degree. This problem is of special concern to people who have occasion to work with old films, but not usually to the general public. Second, cellulose nitrate is a very flammable material that ignites when exposed to even minimal flames or sparks. Most products that contain cellulose nitrate today have been treated to minimize that risk. However, products that were made many years ago do not have the same protection and may pose safety risks to people who come into contact with them.

FOR FURTHER INFORMATION

"Caring for Cellulose Nitrate Film." Conserve O Gram.
http://www.cr.nps.gov/museum/publications/conservogram/14_08.pdf (accessed on September 30, 2005).

"Cellulose Nitrate." Conservation and Art Material Encyclopedia Online.

http://www.mfa.org/_cameo/frontend/material_description.asp?name=cellulose+nitrate&language=1 (accessed on September 30, 2005).

"DiscoverLight™ Membrane." Pierce Biotechnology.

<http://www.piercenet.com/files/5186.pdf> (accessed on September 30, 2005).

Fox, Berry. "Not Fade Away." *New Scientist* (March 1, 2003); 40.

"Nitrocellulose Applications, Characterisation, Production and Storage." Bayer Material Science.

<http://www.azom.com/details.asp?ArticleID=2785> (accessed on September 30, 2005).

See Also Cellulose; Cellulose Acetate; Cellulose Xanthate

OTHER NAMES:	Viscose rayon; xanthate rayon
FORMULA:	$[C_6H_7O_2(OH)_2OCS_2Na]_n$
ELEMENTS:	Carbon, hydrogen, oxygen, sulfur, sodium
COMPOUND TYPE:	Polymer
STATE:	Solid
MOLECULAR WEIGHT:	Very large (50,000 to 100,000 g/mol)
MELTING POINT:	Not applicable
BOILING POINT:	Not applicable
SOLUBILITY:	Soluble in aqueous caustic solutions

KEY FACTS

Cellulose Xanthate

OVERVIEW

Cellulose xanthate (SELL-you-lohs ZAN-thate) is a compound produced from cellulose in the viscose process for making rayon. The subscript “n” in the chemical formula indicates that many molecules chain together to make up the polymer. In this case, between 200 and 400 $C_6H_7O_2(OH)_2OCS_2Na$ molecules combine to make a molecule of cellulose xanthate.

Although rayon is made from a naturally-occurring material, cellulose, it is regarded as a synthetic product. It was the first synthetic fiber made by humans. The method for making rayon was discovered in about 1891 by three English chemists, Charles Frederick Cross (1855-1935), Edward John Bevan (1856-1921), and Clayton Beadle (1868-1917). The discovery occurred while the three chemists were looking for a way of making artificial silk. Although silk is an excellent fiber for making clothes and other materials, it had to be exported from Japan and was, therefore, very expensive.

Interesting Facts

Rayon was first sold in the United States as “artificial silk.” But manufacturers thought the word “artificial” would not appeal to consumers. So they held a contest with a \$1,000 prize to choose a better name for the fabric.

More than 10,000 names were suggested, but the committee appointed to choose a winner rejected all of the names. Finally, a member of the committee suggested the word *rayon*, French for “a ray of light.”

The first plant for the manufacture of rayon in the United States was built in 1910 by the newly-formed Viscose Corporation of America. Sales of the new product were slow at first, but by 1925, rayon had become more popular than silk. Today, the Cross-Bevan-Beadle method of making rayon is only one of four processes used by the industry to make the fiber.

HOW IT IS MADE

The process to make rayon begins with wood pulp or wood chips, usually from spruce or pine trees. The wood is treated with carbon disulfide (CS_2), which converts cellulose in the wood to cellulose xanthate. The purpose of this step is to change an insoluble material (cellulose) into a soluble material (cellulose xanthate). The cellulose xanthate is then dissolved in sodium hydroxide (NaOH), a strong alkali, forming a thick, viscous solution. The viscosity of the solution is responsible for the name given to this method of making rayon, the viscose process.

The cellulose xanthate solution is then allowed to stand for eight to ten hours in a “ripening room.” During this period, the cellulose xanthate slowly gives up its carbon disulfide and reverts to cellulose. The end result of these steps is the recovery of cellulose in a form that is easier to work with and from which long fibers can be made. The fibers are produced by forcing cellulose from the viscous solution through plates called spinnerettes with many small holes in them.

Words to Know

ALKALI A strong base.

AQUEOUS SOLUTION A solution that consists of some material dissolved in water.

CAUSTIC Strongly basic or alkaline, capable of irritating or corroding living tissue.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

VISCOUS Syrupy material that flows slowly.

COMMON USES AND POTENTIAL HAZARDS

Rayon has many properties that make it a desirable fabric. It has a soft, silky feel that is comfortable to the skin. Its fibers are strong, abrasion resistant, easy to dye, and resistant to bleaches and most chemicals. Electric charges do not collect on the fiber, allowing it to remain static-free and to hang smoothly from the body. Rayon does have some disadvantages. It tends to wrinkle easily, shrinks when it gets wet, and is very flammable. Among the many products made from rayon are the following:

- Clothing, such as suits, dresses, sports shirts, ties, lingerie, and work clothes;
- Domestic textiles, such as upholstery, curtains, drapes, bedspreads, sheets, and blankets; and
- Industrial textiles, such as reinforcing threads in rubber products, such as tires, hoses, and conveyor belts; braided cords; and tapes.

Rayon is very flammable, probably its greatest disadvantage in commercial products. During the late 1940s and early 1950s, a number of accidents were reported in which rayon clothing caught fire, resulting in injury or death to wearers. In response to this problem, the U.S. Congress passed the Flammable Fabrics Act of 1953 (FFA), banning the sale of any fabric (such as rayons) likely to catch fire. Since passage of the FFA, manufacturers who use rayon in their products must treat the fiber in some way to prevent it from catching fire.

FOR FURTHER INFORMATION

“Rayon Fiber (Viscose).” Fibersource.

http://www.fibersource.com/f_tutor/rayon.htm (accessed on September 14, 2005).

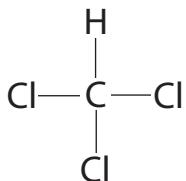
“Rayon Fibers.” Raghavendra R. Hegde, Atul Dahiya, and M. G. Kamath.

<http://www.engr.utk.edu/mse/pages/Textiles/Rayon%20fibers.htm> (accessed on September 14, 2005).

“Rayon Viscose.”

<http://www.swicofil.com/viscose.html> (accessed on September 14, 2005).

See Also Cellulose, Cellulose Nitrate



OTHER NAMES:

Trichloromethane,
trichloroform,
methane trichloride,
methylene trichloride,
methyl trichloride

FORMULA:

CHCl_3

ELEMENTS:

Carbon, hydrogen,
chlorine

COMPOUND TYPE:

Halogenated
hydrocarbon, alkyl
halide (organic)

STATE:

Liquid

MOLECULAR WEIGHT:

119.38 g/mol

MELTING POINT:
 -63.41°C (-82.14°F)

BOILING POINT:
 61.17°C (142.1°F)

SOLUBILITY:

Slightly soluble in
water; soluble in
ethyl alcohol, ether,
acetone, benzene, and
other organic
solvents

K E Y F A C T S

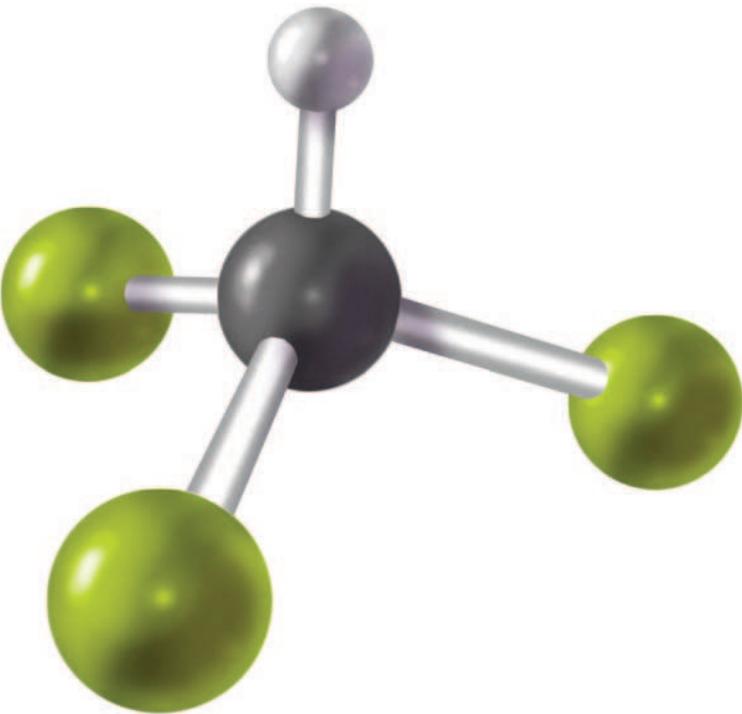
OVERVIEW

Chloroform (KLOR-oh-form) is a clear, colorless, flammable, volatile liquid with a characteristic odor and a sweet taste. It was discovered almost simultaneously in 1831 by American chemist Samuel Guthrie (1782-1848), French chemist Eugene Soubeiran (1797-1858), and German chemist Justus von Liebig (1803-1873). The chemical structure of the compound was determined by French chemist Jean-Baptiste-André Dumas (1800-1884), who suggested its modern name of chloroform in 1834 or 1835. The compound's anesthetic effects on animals were first observed by French physiologist Marie Jean-Pierre Flourens (1794-1867) in 1847.

The potential value of chloroform as an anesthetic for humans was immediately evident. In Great Britain, Liverpool chemist David Walde (1813-1889) suggested to a surgeon friend of his, Sir James Young Simpson (1811-1870), that he might try using chloroform instead of ether to ease the pain of childbirth with his patients. On the evening of November 4, 1847, Simpson and two physician friends experimented

Chloroform. White atom is hydrogen; black atom is carbon; and green atoms are chlorine.

PUBLISHERS RESOURCE GROUP



with chloroform to test its effects and all three passed out. Simpson was convinced of chloroform's value as an anesthetic and began using it in his medical practice. He did this over the objections of many clergymen of the era who believed that women should suffer the pain of childbirth because of Eve's actions in the Garden of Eden, as described in the Bible. The use of chloroform as a painkiller soon spread widely throughout Great Britain and other countries of the world.

In the twenty-first century, chloroform has been almost entirely replaced as an anesthetic by other compounds that are more effective and have fewer undesirable side effects. The compound is now used almost exclusively for the production of hydrochlorofluorocarbons (HCFCs), compounds developed to replace chlorofluorocarbons (CFCs) that have been banned because of their harmful effects on Earth's ozone layer.

HOW IT IS MADE

Chloroform is produced commercially by the reaction between chlorine gas and methane gas. In this reaction, each

Interesting Facts

- The debate over the use of chloroform as an anesthetic during childbirth was largely resolved in 1853 when Queen Victoria (1819–1901) chose to have the procedure during the birth of her son Prince Leopold (1853–1884).
- Chloroform is recommended for the removal of lipstick stains.
- At one time, chloroform was used internally (usually in small doses) to treat a variety of medical problems, including cholera, gonorrhea, colic, cramps, spasms, and convulsions. Improper doses led to problems that were more serious than those being treated, including coma and death.
- Chloroform was once used in a number of consumer and household products, including toothpastes, cough syrups, and skin ointments. Its use for these purposes was banned in the United States in 1976.

of the four hydrogens in methane is replaced, one at a time, by chlorine atoms. The reaction must be carefully controlled, therefore, to ensure that maximum amounts of the desired product (chloroform, in this case) are obtained.

Chloroform can also be produced in the reaction between acetone and bleaching powder (CaOCl_2) in the presence of sulfuric acid (H_2SO_4): $2\text{CH}_3\text{COCH}_3 + 6\text{CaOCl}_2 \rightarrow 2\text{CHCl}_3 + (\text{CH}_3\text{COO})_2\text{Ca} + 3\text{CaCl}_2 + 2\text{Ca(OH)}_2$.

COMMON USES AND POTENTIAL HAZARDS

In the early twenty-first century, more than 90 percent of the chloroform produced in the United States is used in the preparation of hydrochlorofluorocarbons (HCFCs). HCFCs are compounds consisting of carbon, hydrogen, chlorine, and fluorine. They were developed in the 1980s to replace the chlorofluorocarbons (CFCs) that, at the time, were widely used for a number of industrial applications. Replacing CFCs had become necessary because they were also destroying the

Words to Know

ANESTHETIC A substance that is used to temporarily deaden the feeling in a part of the body or to make a person unconscious.

CARCINOGEN A substance that causes cancer in humans or other animals.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

VOLATILE Able to turn to vapor easily at a relatively low temperature.

ozone layer in the stratosphere that protects plants and animals on Earth's surface from harmful ultraviolet radiation.

Chloroform is used to make one specific HCFC known as HCFC-22. HCFC-22 is also known as Freon 22, halocarbon 22, fluorocarbon 22, or R22. By 2003, an estimated 600 million metric tons (660 million short tons) of chloroform were being used in the United States for all commercial purposes. Of the HCFC-22 produced, about 70 percent was being used as a refrigerant, while the remaining 30 percent went for the synthesis of polymers.

In addition to its use in the production of HCFC-22, chloroform has a number of other minor uses, including:

- As a solvent for fats, oils, waxes, rubber, and a number of pharmaceuticals;
- In the manufacture of certain dyes and pesticides;
- As a cleansing agent in various industrial operations;
- In fire extinguishing systems;
- As a glue for methyl methacrylate products; and
- As a laboratory reagent.

Long-term prospects for chloroform are not very promising since the production of HCFCs is scheduled to be phased out over time. Although they have less serious effects on ozone than do the CFCs, they are still regarded as a threat and their use has been restricted under provisions of the Montreal Protocol on Substances that Deplete the Ozone Layer. The U.S. Environmental Protection Agency plans to ban the use of HCFC-22 under the Protocol by 2010.

Chloroform is an irritant to the skin, eyes, and respiratory system. Moderate, short-term exposure may cause redness, pain, and dryness of the skin; redness and pain in the eyes; and coughing, nausea, headache, dizziness, drowsiness, vomiting, abdominal pain, and unconsciousness. At higher doses and longer exposure, chloroform is believed to cause serious damage to the liver and kidneys. It is also thought to be a carcinogen.

FOR FURTHER INFORMATION

“Chloroform.” 11th Report on Carcinogens. National Toxicology Program.

<http://ntp.niehs.nih.gov/ntp/roc/eleventh/profiles/s038chlo.pdf> (accessed on December 29, 2005).

“Chloroform.” International Programme on Chemical Safety.

<http://www.inchem.org/documents/ehc/ehc/ehc163.htm> (accessed on December 29, 2005).

“Occupational Safety and Health Guideline for Chloroform.” Occupational Safety & Health Administration.

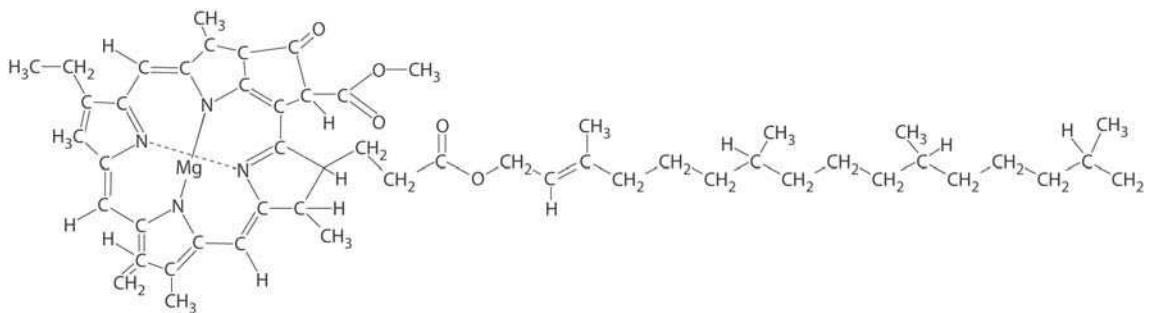
<http://www.osha.gov/SLTC/healthguidelines/chloroform/recognition.html> (accessed on December 29, 2005).

Stratmann, Linda. *Chloroform: The Quest for Oblivion*. Phoenix Mill, UK: Sutton Publishing Co., 2003.

“ToxFAQs™ for Chloroform.” Agency for Toxic Substances and Disease Registry.

<http://www.atsdr.cdc.gov/tfacts6.html> (accessed on December 29, 2005).

See Also Carbon Tetrachloride



FORMULA:

Varies; see Overview.

ELEMENTS:

Carbon, hydrogen, oxygen, nitrogen, magnesium

COMPOUND TYPE:

Organic

STATE:

Solid

MOLECULAR WEIGHT:
608.96–907.47 g/mol

MELTING POINT:
Chlorophyll a: 152.3°C (306.1°F); Chlorophyll b: 125°C (257°F)

BOILING POINT:
Not applicable

SOLUBILITY:
Chlorophyll a and b are insoluble in water, soluble in alcohol, ether, and oils

K E Y F A C T S

OVERVIEW

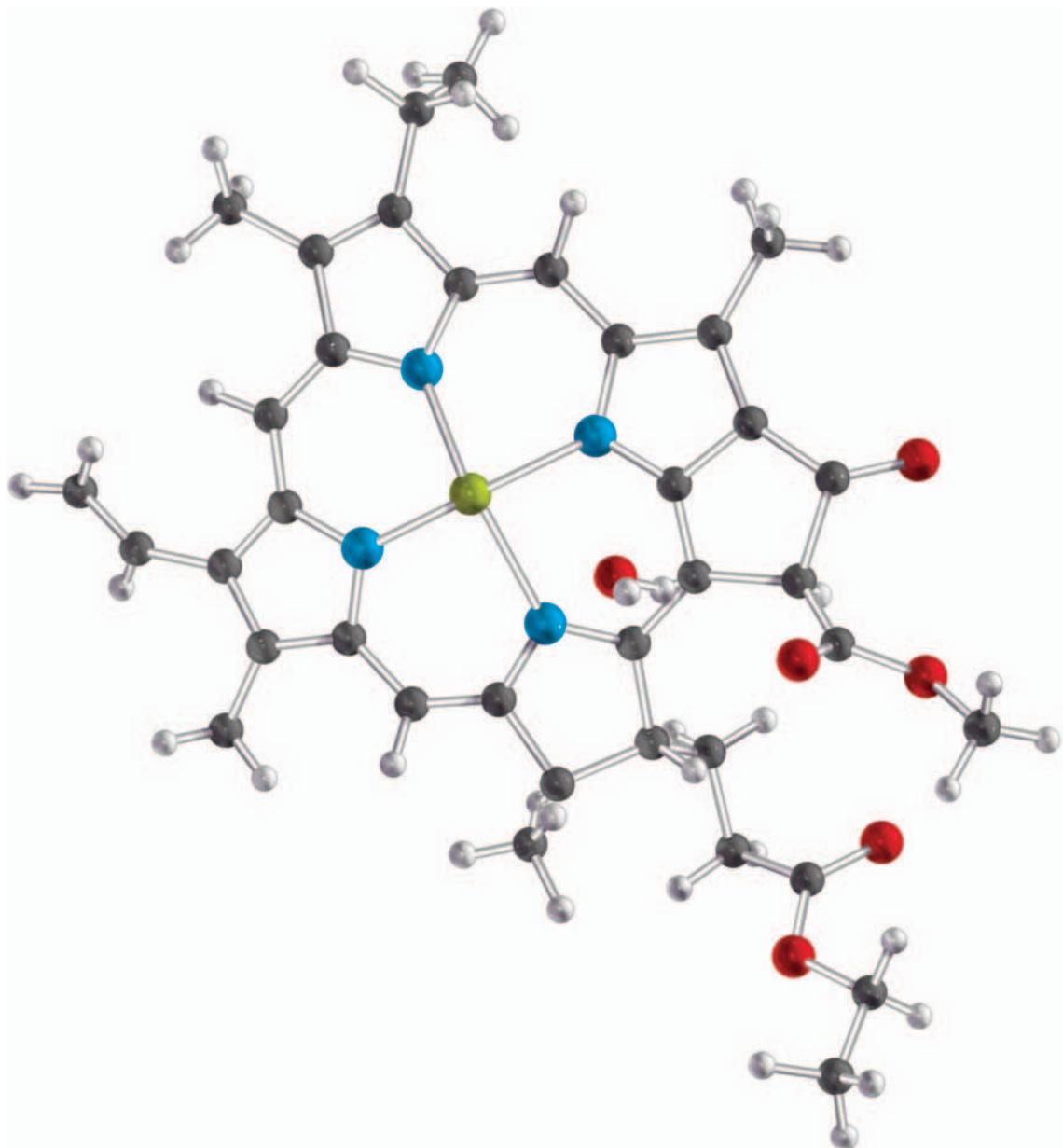
Chlorophyll (KLOR-uh-fill) is the pigment that gives plants, algae, and cyanobacteria their green color. The name comes from a combination of two Greek words, chloros, meaning “green” and phylon, meaning “leaf.” Chlorophyll is the substance that enables plants to create their own food through photosynthesis.

At least five forms of chlorophyll exist. They are:

- chlorophyll a (also known as α -chlorophyll), with a formula of $C_{55}H_{72}O_5N_4Mg$
- chlorophyll b (also known as β -chlorophyll), with a formula of $C_{55}H_{70}O_6N_4Mg$
- Chlorophyll c₁, with a formula of $C_{35}H_{30}O_5N_4Mg$
- Chlorophyll c₂, with a formula of $C_{35}H_{28}O_5N_4Mg$
- Chlorophyll d, with a formula of $C_{54}H_{70}O_6N_4Mg$

Chlorophyll a occurs in all types of plants and in algae. Chlorophyll b is found primarily in land plants. Chlorophyll

CHLOROPHYLL



Chlorophyll. Red atoms are oxygen; white atoms are hydrogen; carbon atoms are black; blue atoms are nitrogen; and green atom is magnesium. PUBLISHERS RESOURCE GROUP

c₁ and chlorophyll c₂ are present in various types of algae. Chlorophyll d is found in red algae.

All forms of chlorophyll have a similar chemical structure. They have a complex system of rings made of carbon and nitrogen known as a chlorin ring. The five forms of

Interesting Facts

- The chemical structure of chlorophyll is very similar to that of hemoglobin, the molecule that transports oxygen in the red blood cells of mammals. The major difference between the two is that hemoglobin contains an atom or iron at the center of a large ring compound, while chlorophyll has an atom of magnesium in the same location.
- Leaves contain compounds called carotenoids that are red, orange, and yellow in color. These colors are masked by the green color of chlorophyll. In fall, plants stop producing chlorophyll, and the red, orange, and yellow of carotenoids become visible. Carotenoids do not perform photosynthesis, although they do transmit light energy to chlorophyll, where photosynthesis takes place.

chlorophyll differ in the chemical groups attached to the chlorin ring. These differences result in slightly different colors of the five chlorophylls.

French chemists Pierre-Joseph Pelletier (1788-1842) and Joseph-Bienaimé Caventou (1795-1877) first isolated chlorophyll in 1817. In 1865, German botanist Julius von Sachs (1832-1897) demonstrated that chlorophyll is responsible for photosynthetic reactions that take place within the cells of leaves. In the early 1900s, Russian chemist Mikhail Tsvett (1872-1920) developed a technique known as chromatography to separate different forms of chlorophyll from each other. In 1929, the German chemist Hans Fischer (1881-1945) determined the complete molecular structure, making possible the first synthesis of the molecule in 1960 by the American chemist Robert Burns Woodward (1917-1979).

HOW IT IS MADE

Plants make chlorophyll in their leaves using materials they have absorbed through their roots and leaves. The synthesis of chlorophyll requires several steps involving

Words to Know

CHLORIN RING A ring of carbon and nitrogen atoms bonded to each other.

CHROMATOGRAPHY Process by which a mixture of substances passes through a column consisting of some material that causes the individual components in the mixture to separate from each other.

PHOTOSYNTHESIS The process by which green plants and some other organisms

using the energy in sunlight to convert carbon dioxide and water into carbohydrates and oxygen.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

complex organic compounds. First, the plant converts a common amino acid, glutamic acid ($\text{COOH}(\text{CH}_2)_2\text{CH}(\text{NH}_2)\text{COOH}$) into an alternative form known as 5-aminolevulinic acid (ALA). Two molecules of ALA are then joined to form a ring compound called porphobilinogen. Next, four molecules of porphobilinogen are joined to form an even larger ring structure with side chains. Oxidation of the larger ring structure introduces double bonds in the molecule, giving it the ability to absorb light energy. Finally, a magnesium atom is introduced into the center of the ring and side chains are added to the ring to give it its final chlorophyll configuration.

COMMON USES AND POTENTIAL HAZARDS

Plants store chlorophyll in their chloroplasts, organelles (small structures) that carry out the steps involved in photosynthesis. Each chloroplast contains many clusters of several hundred chlorophyll molecules called photosynthetic units. When a photosynthetic unit absorbs light energy, chlorophyll molecules move to a higher energy state, initiating the process of photosynthesis. The overall equation for the process of photosynthesis is $6\text{CO}_2 + 6\text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2$.

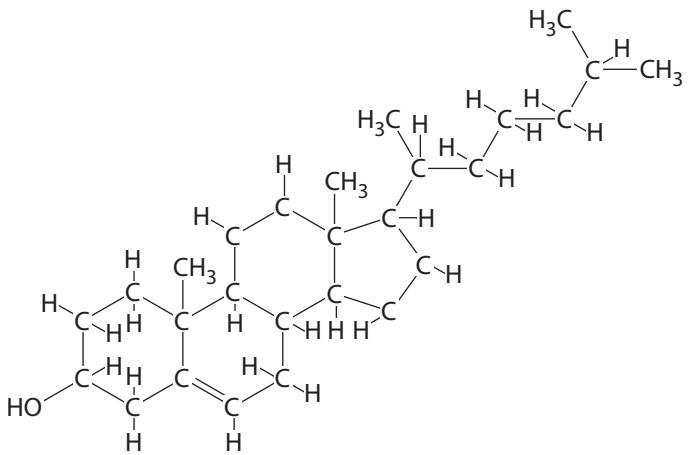
That simple equation does not begin to suggest the complex nature of what happens during photosynthesis. Botanists divide that process into two major series of reactions: the light reactions and the dark reactions. In the light reactions, plants

use the energy obtained from sunlight to make two compounds, adenosine triphosphate (ATP) and nicotinamide adenine dinucleotide phosphate (NADPH). ATP and NADPH are not themselves components of carbohydrates, the final products of photosynthesis. Instead, they store energy that is used to make possible a series of thirteen different chemical reactions that occur during the dark stage of photosynthesis that result in the conversion of carbon dioxide and water to the simple carbohydrate glucose ($C_6H_{12}O_6$).

FOR FURTHER INFORMATION

- Attenborough, David. *The Private Life of Plants*. Princeton, NJ: Princeton University Press, 1995.
- Buchanan, B. B., W. Gruissem, and R. L. Jones. *Biochemistry and Molecular Biology of Plants*. Rockville, MD: American Society of Plant Physiologists, 2000.
- "Chlorophyll and Chlorophyllin." The Linus Pauling Institute Micronutrient Information Center.
<http://lpi.oregonstate.edu/infocenter/phytochemicals/chlorophylls/> (accessed on October 3, 2005).
- May, Paul. "Chlorophyll." School of Chemistry, University of Bristol.
http://www.chm.bris.ac.uk/motm/chlorophyll/chlorophyll_h.htm (accessed on October 3, 2005).
- Steer, James. "Structure and Reactions of Chlorophyll."
<http://www.ch.ic.ac.uk/local/projects/steer/chloro.htm> (accessed on October 3, 2005).

See Also Carbon Dioxide; Water



OTHER NAMES:
See Overview.

FORMULA:
 $C_{27}H_{45}OH$

ELEMENTS:
Carbon, hydrogen,
oxygen

COMPOUND TYPE:
Sterol

STATE:
Solid

MOLECULAR WEIGHT:
386.65 g/mol

MELTING POINT:
148.5°C (299.3°F)

BOILING POINT:
Not applicable;
decomposes at about
360°C (680°F)

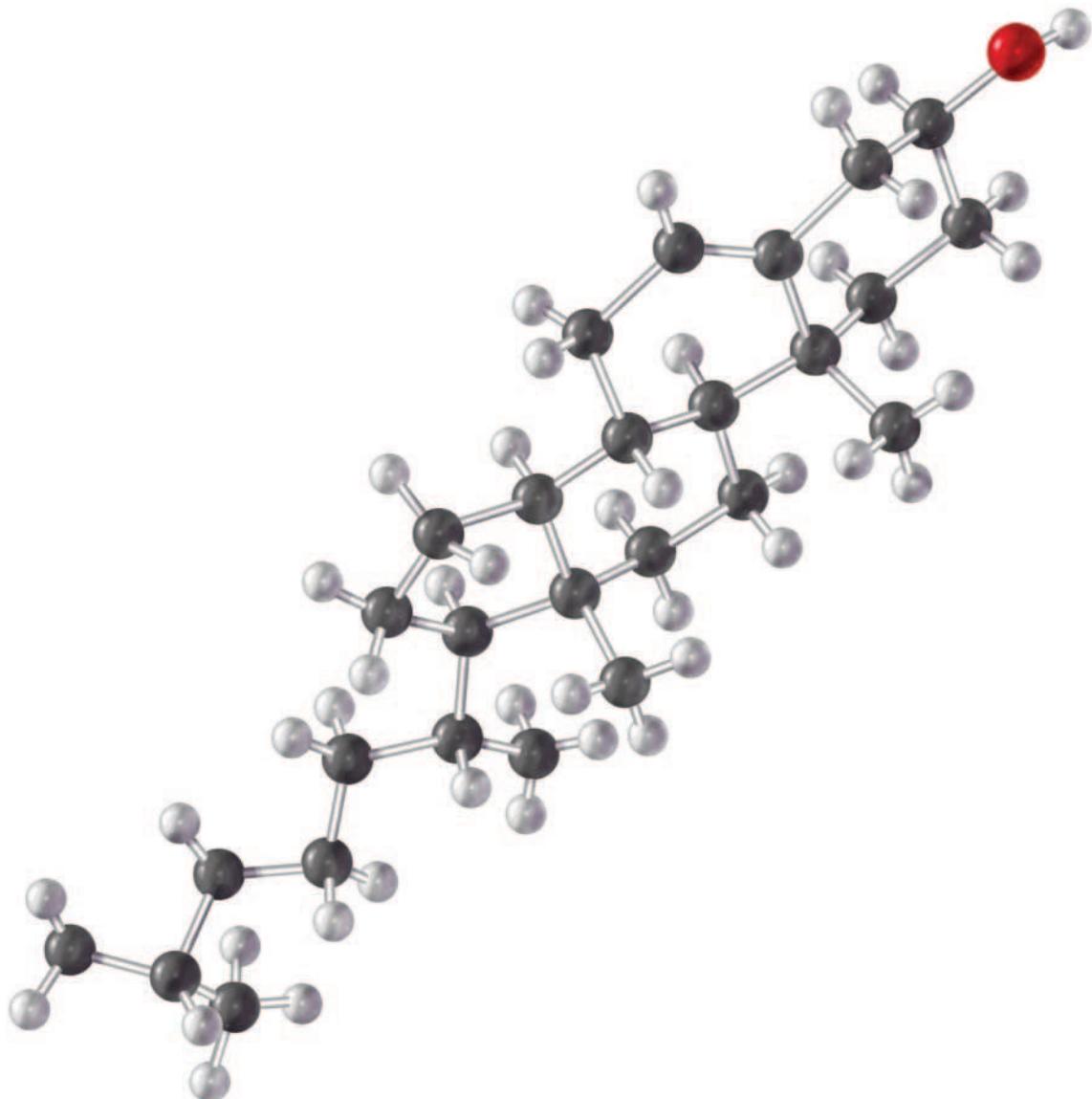
SOLUBILITY:
Very slightly soluble
in water; soluble in
acetone, alcohol,
ether, benzene, and
other organic
solvents.

KEY FACTS

OVERVIEW

Cholesterol (koh-LESS-ter-ol) is also known as cholesterin; cholest-5-3n-3β-ol; 5-cholest-3-β-ol; 3β-hydroxy-5-cholestene; and 10,13-dimethyl-17-(6-methylheptan-2-yl)-2,3,4,7,8,9,11,12,14,15,16,17-dodecahydro-1H-cyclopenta[*a*]phenanthren-3-ol. It is a waxy white or pale yellow solid with virtually no taste or odor. It is present in the bodies of all higher animals, especially in the brain and spinal cord. Chemically, cholesterol is classified as a fat, a member of the lipid family. Fats are the product of the trihydric alcohol (alcohol with three -OH groups) glycerol and a fatty acid. Fatty acids are organic acids with many carbon atoms, usually eight or more.

The first description of cholesterol was written by a French scientist by the name of Poulletier de la Salle (dates not available), who isolated the compound from gallstones, a rich source of the substance. In 1816, de la Salle's compatriot Michel Eugène Chevreul (1786–1889) suggested the name of cholesterine for cholesterol, combining the French words for



Cholesterol. Red atom is oxygen; white atoms are hydrogen; and black atoms are carbon. PUBLISHERS RESOURCE GROUP

bile (chole) and stone (stereos). More than a century passed before an accurate understanding of the chemical structure and behavior of cholesterol was attained. Finally, during the 1910s, research by the German chemists Heinrich Otto Wieland (1877–1957) and Adolf Windaus (1876–1959) solved most of the fundamental problems about cholesterol, an accomplishment for which the two men were awarded the

Interesting Facts

- The amount of cholesterol in a person's body depends essentially on two factors: genetics and diet. The ability of a person's body to produce cholesterol is determined by his or her genetic component. Some people simply make high levels of cholesterol, and other produce lower levels. Those levels can be affected to some extent by the kinds and amounts of cholesterol-containing foods (such as dairy products and organ meats) one includes in his or her daily diet.

1927 (Wieland) and the 1928 (Windhaus) Nobel Prizes in chemistry. All that remained was the determination of the structural formula for cholesterol and the first synthesis of the compound, both achieved by the American chemist Robert Burns Woodward (1917-1979) in 1955.

HOW IT IS MADE

The primary source of cholesterol is the liver. After cholesterol is produced, the liver packages cholesterol molecules with proteins in units known as lipoproteins, which are then distributed throughout the body by way of the bloodstream. Some cholesterol is deposited in cells, where it is used as the raw material in the synthesis of a number of biologically essential compounds, such as vitamin D₃, steroids (such as estrogen and testosterone), and the bile acids, used by the body to digest foods. Excess cholesterol not needed by cells remains in the bloodstream until it is returned to the liver.

Commonly, the body produces more cholesterol than cells need. Or a person takes in cholesterol in his or her diet in excess of that needed by the body. In such cases, cholesterol may accumulate in the bloodstream and, at some point, be deposited on the walls of blood vessels. As these deposits grow larger, they may impeded the flow of blood, eventually leading to heart problems or stroke.

Words to Know

FAT An ester formed in the reaction between glycerol ($C_3H_5(OH)_3$) and a fatty acid, an organic acid with more than eight carbon atoms.

LIPIDS Organic compounds that are insoluble in water, but soluble in most organic solvents, such as alcohol, ether, and acetone.

PRECURSOR A compound that gives rise to some other compound in a series of reactions.

STEROL Organic compound containing one or more ring systems and a hydroxy (-OH) group

Lipoproteins produced in the liver take one of two forms. Low-density lipoproteins (LDL) carry cholesterol directly to cells. Because this cholesterol serves a useful biological function, it is sometimes called “good” cholesterol. High-density lipoproteins (HDL), by contrast, bypass cells and remain in the blood stream. The cholesterol in HDL lipoproteins serves no useful function in the body and is often referred to as “bad” cholesterol. The role of cholesterol in the human body and dietary methods of maintaining the correct HDL:LDL ratio in the blood has been the subject of some controversy and considerable educational programs in recent decades.

A relatively small amount of cholesterol is used in making some drugs and cosmetics. The cholesterol required for these commercial purposes is obtained from natural sources, primarily spinal fluid taken from cattle who have been destroyed for meat products.

COMMON USES AND POTENTIAL HAZARDS

Cholesterol plays three primary roles in the human body. First, it helps the body digest fatty foods. During digestion, the liver converts cholesterol into a yellowish-green liquid called bile. Bile helps break down fatty foods into a form that can be absorbed by the body. Second, cholesterol is a precursor to a number of important hormones, including androgen, estrogen, progesterone, and testosterone. Third, cholesterol is also a precursor to vitamin D, which has a number of essential functions, including the maintenance of healthy

bones and nervous system, proper growth, muscle tone, production of insulin, normal reproductive functions, and proper immune system function.

FOR FURTHER INFORMATION

“Cholesterol.” Chemical Land 21.

<http://www.chemicaland21.com/arokorhi/lifescience/phar/CHOLESTEROL.htm> (accessed on October 3, 2005).

“Cholesterol.” Medline Plus.

<http://www.nlm.nih.gov/medlineplus/cholesterol.html> (accessed on October 3, 2005).

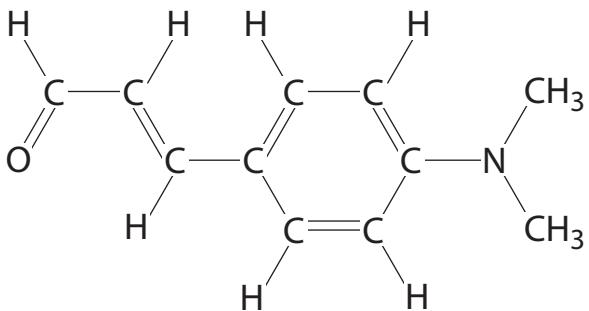
“Cholesterol, Other Lipids, and Lipoproteins.” In *In Depth Report*. Edited by Julia Goldrosen. Atlanta: A.D.A.M., 2004.

Goldstein, Joseph P., and Michael S. Brown. “Cholesterol: A Century of Research.” *HHMI Bulletin*, 2003.

<http://www.hhmi.org/bulletin/sept2003/cholesterol/century.html> (accessed on October 3, 2005).

“The Nobel Prize in Physiology or Medicine 1985.” Nobelprize.org.
<http://nobelprize.org/medicine/laureates/1985/press.html> (accessed on October 3, 2005).

“The Overall Synthesis of Cholesterol.” Synthesis of Cholesterol
<http://www.chm.bris.ac.uk/webprojects2000/ahester/synthchol.html> (accessed on October 3, 2005).



OTHER NAMES:

See Overview.

FORMULA:
 $C_6H_5CH=CHCHO$

ELEMENTS:
Carbon, hydrogen,
oxygen

COMPOUND TYPE:
Aromatic aldehyde
(organic)

STATE:
Liquid

MOLECULAR WEIGHT:
132.16 g/mol

MELTING POINT:
−7.5°C (18°F)

BOILING POINT:
246°C (475°F)

SOLUBILITY:
Very slightly soluble
in water; soluble in
alcohol, ether, and
chloroform

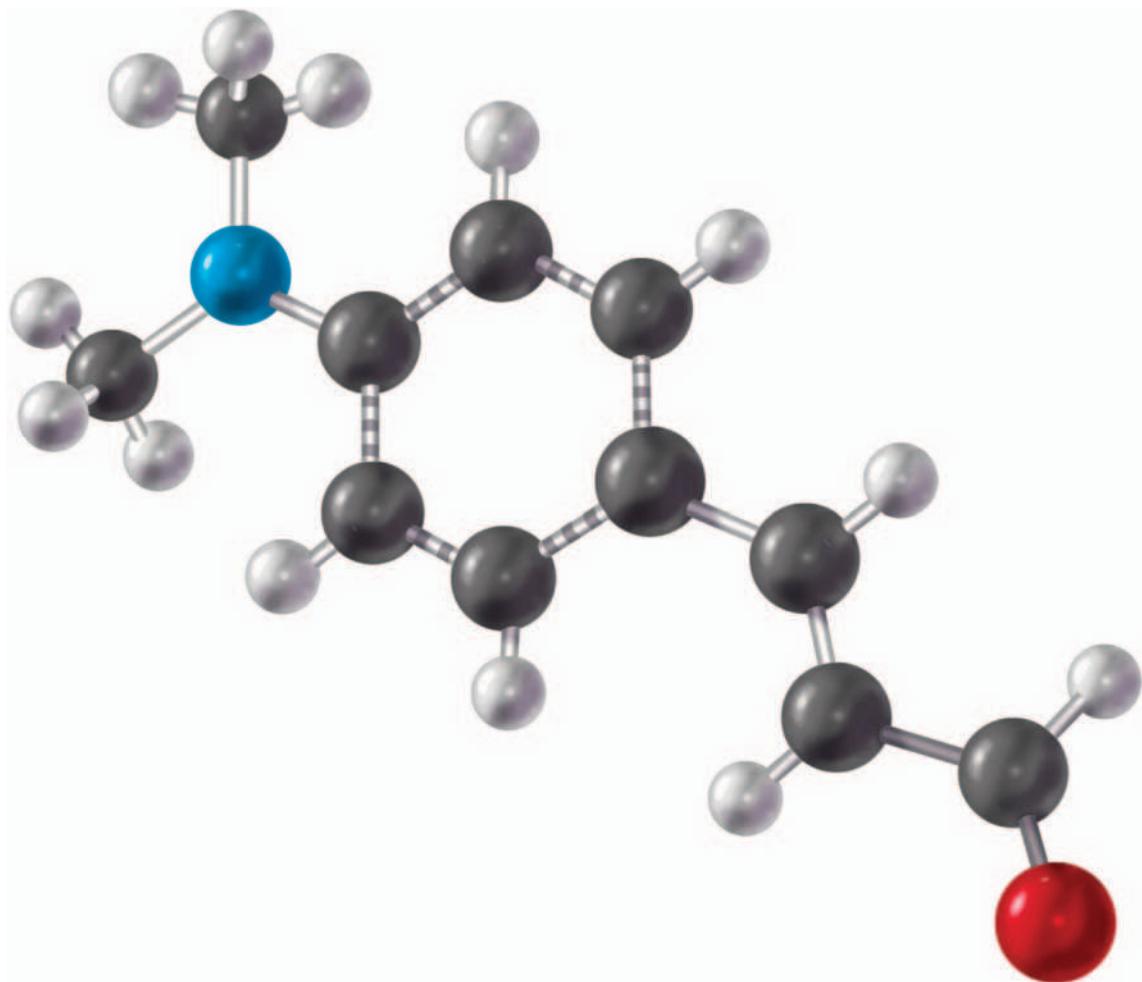
**K
E
Y
F
A
C
T
S**

OVERVIEW

Cinnamaldehyde (sin-uh-MAL-duh-hide) is also known as cinnamic aldehyde; 3-phenyl-2-propenal; cinnamyl aldehyde; phenylalacrolein; cinnamal; and trans-cinnamaldehyde. It is a yellowish oily, liquid with a sweet taste and a cinnamish odor responsible for the characteristic taste and odor of cinnamon spice. It occurs naturally in the bark of the cinnamon tree, *Cinnamomum zeylanicum*, which is native to Sri Lanka and India, and has been cultivated in other parts of the world, such as Brazil, Jamaica, and Mauritius. Cinnamaldehyde is also found in other members of the *Cinnamomum* species, including cassia and camphor.

The molecular formula for cinnamaldehyde was determined in 1834 by the French chemists Jean Baptiste André Dumas (1800–1884) and Eugène Melchior Péligot (1811–1890), although its structural formula was deciphered only in 1866 by the German chemist Emil Erlenmeyer (1825–1909).

Primary uses for Cinnamaldehyde are as a food flavoring and a medical herb. It has been used for many centuries to



Cinnamaldehyde. Red atom is oxygen; white atoms are hydrogen; black atoms are carbon; and blue atom is nitrogen. Gray sticks show double bonds; Striped sticks show a benzene ring.

PUBLISHERS RESOURCE
GROUP

treat a wide variety of disorders, ranging from the common cold and the flu to diarrhea to cancer. It was mentioned in Chinese medical texts as early as 2700 BCE and was described in a famous Chinese medical text, the *Tang Materia Medica*, written in 659 CE. It has also been used for centuries in Ayurvedic medicine, the ancient healing art of India.

HOW IT IS MADE

Cinnamaldehyde is prepared commercially by treating the bark of the *Cinnamomum zeylanicum* tree with steam. The aldehyde dissolves in the steam and can then be

Interesting Facts

- Cinnamaldehyde is a component of sprays used as cat and dog repellents.

extracted as the steam cools and condenses to form cold water, in which the compound is much less soluble. Cinnamaldehyde can also be synthesized by reacting benzaldehyde (C_6H_5CHO) with acetaldehyde (CH_3CHO). The two compounds condense with the elimination of water to form cinnamaldehyde.

COMMON USES AND POTENTIAL HAZARDS

In addition to its uses as a herbal remedy, the primary use of cinnamaldehyde is as a food additive to enhance the flavor and/or odor of food products. It is used most commonly in cake mixes, chewing gums, chocolate products, synthetic cinnamon oils, cola drinks, ice creams, soft drinks, and vermouth. The compound is also added to a number of cosmetics and home care products to improve their odor. Such products include deodorants, detergents, mouthwashes, perfumes, sanitary napkins, soaps, and toothpastes. Finally, cinnamaldehyde is used to some extent in agriculture as an insecticide and fungicide.

No safety concerns about the use of cinnamaldehyde as an herbal remedy, food additive, or pesticide has been expressed. The U.S. Food and Drug Administration (FDA) has classified the compound as a generally-recognized as safe (GRAS) food additive, permitting its continued use as a food additive in the United States.

Words to Know

AROMATIC COMPOUND A compound whose chemical structure is based on that of benzene (C_6H_6).

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

FOR FURTHER INFORMATION

“Cinnamaldehyde.” Chemical Land 21.

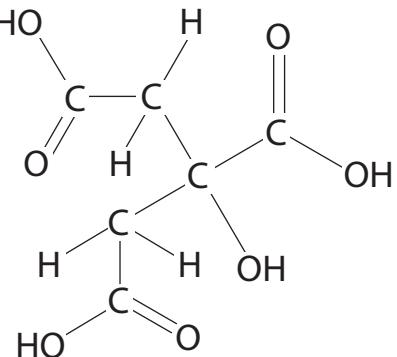
<http://www.chemicaland21.com/arokorhi/specialtychem/perchem/CINNAMALDEHYDE.htm> (accessed on October 3, 2005).

“Cinnamaldehyde (o40506) Fact Sheet.” U.S. Environmental Protection Agency.

http://www.epa.gov/pesticides/biopesticides/ingredients/factsheets/factsheet_o40506.htm (accessed on October 3, 2005).

“Oleum Cinnamomi (U. S. P.) Oil of Cinnamon.” King’s American Dispensatory (1898).

http://www.ibiblio.org/herbmed/eclectic/kings/cinnamomum_oleu.html (accessed on October 3, 2005).



OTHER NAMES:

See Overview.

FORMULA:

$\text{HOOCCH}_2\text{C(OH)(COOH)CH}_2\text{COOH}$

ELEMENTS:

Carbon, hydrogen, oxygen

COMPOUND TYPE:

Carboxylic acid (organic)

STATE:

Solid

MOLECULAR WEIGHT:

192.12 g/mol

MELTING POINT:

153°C (307°F)

BOILING POINT:

Not applicable; decomposes above 175°C (347°F)

SOLUBILITY:

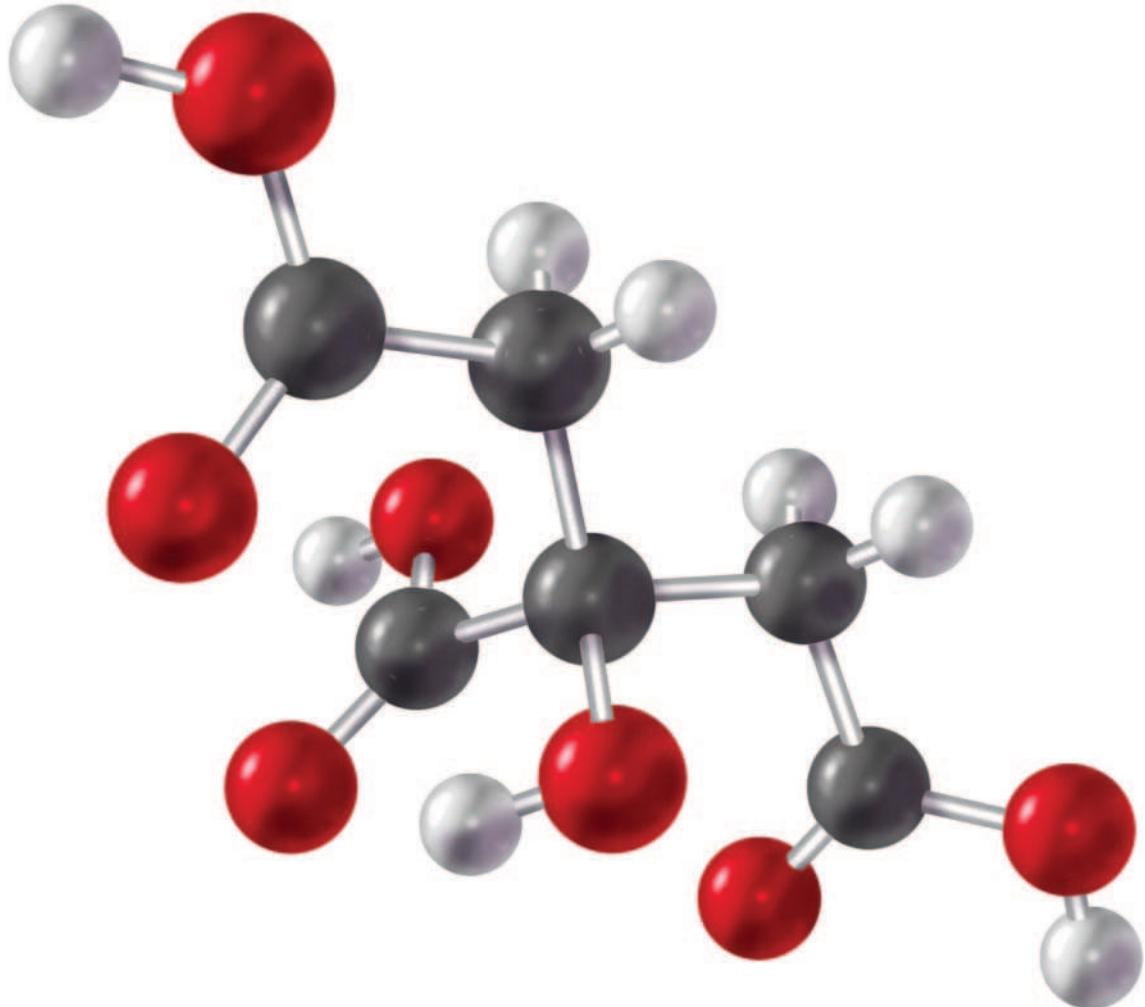
Soluble in water, alcohol, ether, and other organic solvents

KEY FACTS

OVERVIEW

Citric acid (SIT-rik AS-id) is also known as 2-hydroxy-1,2,3-propanetricarboxylic acid and β -hydroxytricarballylic acid. It is a common constituent of plant and animal tissues. Its presence is especially noticeable in citrus fruits, such as lemons, limes, oranges, tangerines, grapefruits, and kumquat, which get their name from the acid. In pure form, citric acid is a colorless, translucent, odorless crystalline or powdery material with a pleasantly acidic taste. It frequently occurs as the monohydrate, with a single molecule of water associated with each citric acid molecule. The formula for the monohydrate is $\text{HOOCCH}_2\text{C(OH)(COOH)CH}_2\text{COOH}\cdot\text{H}_2\text{O}$. The monohydrate is efflorescent, meaning that it tends to lose its water of hydration when exposed to the air.

Citric acid plays an important role in metabolism, the set of chemical reactions that occur when cells break down fats, carbohydrates, and other compounds to produce energy and compounds needed to build new cells and tissues. In fact, the series of reactions by which carbohydrates are



Citric acid. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. PUBLISHERS RESOURCE GROUP

converted to energy is generally known as the citric acid cycle because of the fundamental role played by the compound in those reactions. The citric acid cycle is also known as the Krebs cycle, after the German-British biochemist Sir Hans Adolf Krebs (1900–1981), who discovered the series of reactions in 1937. Citric acid also acts as an antioxidant, a substance that rid the body of molecules called free radicals that can damage healthy cells, promote cancer, and bring about ageing.

The discovery of citric acid is often credited to the Arab alchemist Jabir Ibn Hayyan (721–815), also known by his

Interesting Facts

- The addition of citric acid to candy gives the product a “super sour” taste.

Latin name of Geber. Geber described a substance with all the properties that we equate with citric acid today, but he knew nothing about its chemical structure. The first person to isolate the compound as a pure substance was the Swedish chemist Karl Wilhelm Scheele (1742–1786), who obtained citric acid from the juice of lemons. By the mid-nineteenth century, citric acid was being produced commercially in Italy from lemons and other citrus fruits.

An important step in the commercial manufacture of citric acid occurred in 1892 when the German microbiologist Carl Wehmer (dates not available) found that citric acid could be produced by the penicillin mold. Wehmer's discovery paved the way for large-scale industrial production of citric acid. In 1917, the American chemist James Currie (dates not available) made another important breakthrough in the synthesis of citric acid. While studying the process of fermentation in cheese making, he discovered that the mold *Aspergillus niger* is able to convert sugar to citric acid. After Currie joined the pharmaceutical company Pfizer in 1917, he developed a process called SUCIAC—sugar under conversion into citric acid—by which the compound could be made in mass quantities. That process eventually became the primary method by which citric acid is produced today.

HOW IT IS MADE

At one time, citric acid was obtained primarily from citrus fruits, such as lemons and limes. Today, it is produced synthetically using the *Aspergillus niger* mold, as described above. The citric acid produced in this reaction is purified by crystallization. The anhydrous form crystallizes from hot water, and the monohydrate form from cold water.

Words to Know

ALCHEMY An ancient field of study from which the modern science of chemistry evolved.

FERMENTATION Chemical process by which yeasts or molds break down sugar into alcohol and carbon dioxide.

METABOLISM Process including all of the chemical reactions that occur in cells by which fats, carbohydrates, and other compounds are broken down to produce

energy and the compounds needed to build new cells and tissues.

MORDANT Substance used in dyeing and printing that reacts chemically with both a dye and the material being dyed to help hold the dye permanently to the material.

POLYMER Compound consisting of very large molecules made of one or two small repeated units called monomers.

COMMON USES AND POTENTIAL HAZARDS

Citric acid is added to foods, drinks, and medicines to make them more acidic. Increasing the acidity of these products not only gives them a tart taste, but also prevents the growth of bacteria. Citric acid is also used to preserve the flavors of canned fruits and vegetables and to maintain the proper acidic level of jams and jellies that will help them gel.

In addition to its use as a food additive, citric acid has a number of other commercial and industrial applications, including the following:

- As a sequestering agent to remove small amounts of metals in a solution. A sequestering agent is a substance that surrounds and captures some other substance (such as metals) and removes them from a solution;
- In the cleaning and polishing of stainless steel and other metals;
- As a mordant in the dyeing of cloth;
- In the production of certain kinds of polymers;
- For the removal of sulfur dioxide from the waste gases produced at smelters; and
- As a builder in detergents, a substance that increases the detergent's cleaning efficiency, usually by maintaining

the proper acidity or softening the water in which the detergent acts.

FOR FURTHER INFORMATION

“Citric Acid.” Jungbunzlauer.

http://www.jungbunzlauer.com/products/product_1.html
(accessed on October 3, 2005).

“Citric Acid, Anhydrous.” J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/c4735.htm>
(accessed on October 3, 2005).

Kuntz, Lynn A. “Acid Basics.” *Food Product Design* (May 1993).

Also available online at

http://www.foodproductdesign.com/toolbar_library.html
(accessed on October 3, 2005).

OTHER NAMES:
Not applicable

FORMULA:
Not applicable

ELEMENTS:
Carbon, hydrogen,
oxygen, nitrogen

COMPOUND TYPE:
Organic

STATE:
Solid

MOLECULAR WEIGHT:
100,000 to 150,000
g/mol

MELTING POINT:
Not applicable

BOILING POINT:
Not applicable

SOLUBILITY:
Not applicable

KEY FACTS

OVERVIEW

Collagen (KO-lah-jen) is the most abundant protein in the animal kingdom. Approximately one third of the protein in a mammal's body is collagen. It makes up a major portion of the connective tissue found in skin, joints, ligaments, muscles, tendons, and bones. Different types of collagen occur in different species. That is, the collagen found in humans is somewhat different from that found in cows, and both human and cow collagen differ from the collagen found in other animals, such as dogs and cats. Nonetheless, all forms of collagen have a common molecular structure that consists of three rope-like strands intertwined with each other. Each strand (called tropocollagen) is a polymer consisting of amino acids, the most common of which are glycine, proline, and hydroxyproline. Collagens differ from each other in regard to the relative amounts of each amino acid present.

HOW IT IS MADE

Collagen is synthesized in the bodies of all higher vertebrates in a complex series of reactions that begin with the

linking of amino acids with each other. Amino acids are organic compounds that contain both the carboxyl group (-COOH) and the amino group (-NH₂). The amino acids are joined to each other in very long chains known as polypeptides (a word meaning “many amino acids”). After assembly, polypeptide chains intertwine with each other in groups of three to make large molecules called procollagen (“early collagen”), which is then cut into smaller pieces by enzymes designed especially for that purpose. The three strands that make up each collagen group are held together by the action of vitamin C molecules.

COMMON USES AND POTENTIAL HAZARDS

All of the connective tissue in the body contains collagen. This collagen is flexible but not stretchy and will tear if pulled too strongly. Collagen is also a primary component of cartilage, the material that fills the gaps between bones and fills out structures that are firm but somewhat soft, such as the end of the nose and ears. There are at least a dozen types of collagen in the body, each one with a slightly different physical and chemical structure.

Type I is the most abundant kind of collagen. It forms long strands that criss-cross the spaces between cells and is found throughout the body, particularly in tendons, bones, and scars. Type II collagen is found in cartilage, and Type III collagen in granulation tissue that forms when wounds heal. Type IV collagen comprises the basal lumina, the membrane that supports skin and internal surfaces. The eight other types of collagen all have their specific functions in the body.

Plastic surgeons use collagen to fill wrinkles and lines in their patients’ faces and bodies. As people age, the tissue that fills out their faces and makes the skin look smooth and tight breaks down, leaving lines and wrinkles. Injecting collagen into wrinkles fills the area under the skin with tissue that puffs out the skin, making it firm and smooth again. Doctors also use collagen to fill in scars. In most cases, collagen treatments are not permanent. The collagen breaks down or is absorbed and additional treatments are necessary to remove wrinkles.

Interesting Facts

- When collagen is treated with boiling water, it is converted to gelatin, a widely used household product. Popular desserts such as Jell-O® are made from collagen taken from the hooves, bones, and connective tissue of cows and pigs.
- In the cooking method known as “braising”, meats are heated for a long time over low heat. During braising, collagen molecules “melt” (break apart) producing a tender meaty product in a rich broth that turns to gelatin as it cools. By contrast, cooking meat over high heat causes proteins to shrink and become tougher.

Several health disorders result from inadequate or abnormal collagen in the body. Scurvy, caused by a deficiency of vitamin C, is the best known of these. When adequate amounts of vitamin C are not present in the body, collagen molecules break apart and muscles and joints are damaged. Ehlers-Danlos syndrome is a disorder that occurs when collagen molecules are weakened, causing bruising of the skin, rupture of arteries or intestines, overly flexible joints, skin and bone deformities, and hip dislocations. Osteogenesis imperfecta results in poorly formed bones and causes the sclera (the white part of the eye) to turn blue.

Words to Know

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

POLYPEPTIDE A molecule composed of many amino acids in a chain.

FOR FURTHER INFORMATION

“Collagenase and Collagen: What Is Collagen?” BioSpecifics Technologies Corporation.

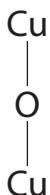
<http://www.biospecifics.com/collagendefined.html> (accessed on December 6, 2005).

“Collagens.” Kimball’s Biology Pages.

<http://users.rcn.com/jkimball.ma.ultranet/BiologyPages/C/Collagens.html> (accessed on December 6, 2005).

Goodsell, David S. “Collagen: Your Most Plentiful Protein.” Protein Data Bank.

http://www.rcsb.org/pdb/molecules/pdb4_1.html (accessed on December 6, 2005).



OTHER NAMES:
See Overview.

FORMULA:
 Cu_2O

ELEMENTS:
Copper, oxygen

COMPOUND TYPE:
Metallic oxide

STATE:
Solid

MOLECULAR WEIGHT:
143.09 g/mol

MELTING POINT:
1235°C (2255°F)

BOILING POINT:
Decomposes at
1800°C (3300°F)

SOLUBILITY:
Insoluble in water;
soluble in inorganic
acids, such as hydro-
chloric, sulfuric, and
nitric acids

KEY FACTS

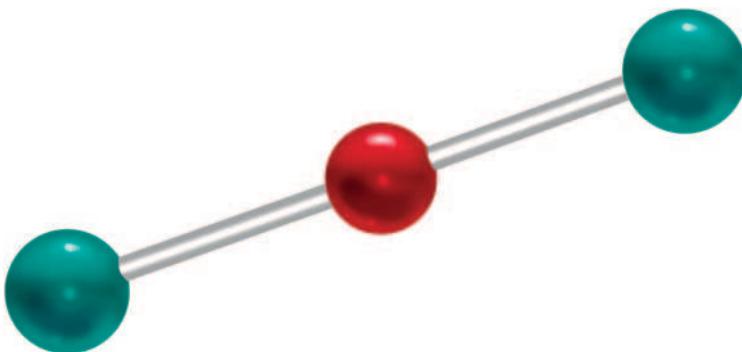
OVERVIEW

Copper(I) oxide (KOPP-er one OK-side) is also known as cuprous oxide, red copper oxide, copper protoxide, copper hemioxide, and copper suboxide. It is a yellowish, red, or brown crystalline substance, depending on its method of preparation. It does not burn and is stable in dry air. In moist air, it slowly changes to copper(II) oxide (CuO). The compound has been used by humans for thousands of years, first as a pigment in glazes, and later in fungicides, electronic components, and industrial reactions.

In 1883, copper(I) oxide was the first substance found to have semiconducting properties. A semiconductor conducts an electric current, although not nearly as efficiently as a conductor like copper, gold, or silver. Semiconductor components are now widely used in computer chips, although they are now made from silicon rather than copper(I) oxide.

HOW IT IS MADE

Copper(I) oxide occurs naturally as the mineral cuprite. It forms when native copper (copper as an element) is exposed



Copper(I) oxide. Red atom is oxygen and turquoise atoms are copper. PUBLISHERS RESOURCE GROUP

to the air, although the process occurs very slowly. Copper(I) oxide can also be made synthetically by exposing copper metal to a high concentration of oxygen at an elevated temperature. Care must be taken, however to prevent the copper from oxidizing completely to form copper(II) oxide. A number of other industrial methods are also available for the synthesis of copper(I) oxide. For example, electrolysis of a solution of sodium chloride using copper electrodes results in the formation of copper(I) oxide.

COMMON USES AND POTENTIAL HAZARDS

Copper(I) oxide is used as a pigment in porcelain glazes and stained glass. In opaque glass, it provides a bright brick-red color if large enough crystals are used. Smaller crystals give a yellowish color. The use of copper(I) oxide as a pigment in glazes dates back to the time of ancient Egypt.

Many antifouling marine paints contain copper(I) oxide. Antifouling paints are paints that prevent the formation of barnacles and other organisms on the bottom of a boat. The compound is also used as an antifungal agent, a substance that kills mildew, rust, and other types of fungus. Fungicides based on copper(I) oxide are commonly used on a variety of crops susceptible to attack by such organisms. Copper(I) oxide acts by inhibiting the growth of fungal spores (from which new plants develop) rather than killing mature fungi.

Copper(I) oxide is used in photoelectric cells, which are cells that generate an electric current when exposed to light.

Interesting Facts

- Copper metal is sometimes used for roofing on notable buildings, such as art museums or state capitols. Over time, the copper metal develops a patina, a thin layer of color that adds to the elegance of the roof. The patina consists primarily of copper(I) oxide.
- In 1904, the German physicist Wilhelm Hallwachs (1859–1922) discovered that a combination of copper metal and copper(I) oxide displays the photoelectric effect, in which a beam of light causes electricity to flow through a material.

Photoelectric materials are usually semiconductors. Copper(I) oxide photoelectric cells do not generate much power, but they react rapidly to changes in light levels. This property makes them useful as light detectors, which have many applications from cameras with automatic adjustment settings to automatic door openers and burglar alarms.

In the 1980s, a ceramic form of copper(I) oxide was found to have superconducting properties at temperatures higher than previously known superconductors. Superconductors have the ability to carry an electric current virtually without resistance. Once a current is initiated in a superconductor, it continues to travel through the material essentially forever. Superconductor research may lead to new technologies, from cheaper electrical power to magnetically levitated high-speed trains.

Exposure to moderate amounts of copper(I) oxide can be irritating to the skin, eyes, and respiratory system. If heated, it gives off copper fumes that can cause symptoms similar to the common cold and may discolor the skin and hair. If swallowed, copper(I) oxide has a metallic taste and may induce vomiting, nausea, and stomach pain. In the most severe cases, copper(I) oxide may cause black or tarry stools, jaundice, and bloody vomiting, with long-term damage to the liver. In the environment, copper(I) oxide is very toxic to fish and crustaceans.

Words to know

ELECTROLYSIS Process in which an electric current is used to bring about chemical changes.

OXIDIZE To combine with oxygen.

PHOTOELECTRIC CELL A device that generates electricity when exposed to light.

SEMICONDUCTOR A substance that allows electricity to flow through it at a lesser

rate than a conductor, though more than a non-conductive substance.

SUPERCONDUCTOR A substance that allows electricity to flow through it with very little resistance.

SYNTHESIS Chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

FOR FURTHER INFORMATION

“Cuprite.” Dale Minerals International.

<http://www.webmineral.com/data/Cuprite.shtml> (accessed on October 3, 2005).

“Cuprous oxide.” Hummel Croton, Inc.

http://www.hummelcroton.com/msds/cu2o_m.html (accessed on October 3, 2005).

Patnaik, Pradyot. *Handbook of Inorganic Chemicals*. New York: McGraw Hil, 2003, 271–273.

Richardson, H. W., ed. *Handbook of Copper Compounds and Applications*. New York: Marcel Dekker, 1997.

See Also Copper(II) Oxide



OTHER NAMES:

Cupric oxide; copper monoxide; black copper oxide

FORMULA:

CuO

ELEMENTS:

Copper, oxygen

COMPOUND TYPE:

Metallic oxide

STATE:

Solid

MOLECULAR WEIGHT:

79.54 g/mol

MELTING POINT:

1,446°C (2,670°F)

BOILING POINT:

Not applicable;
decomposes

SOLUBILITY:

Insoluble in water and
organic solvents;
soluble in dilute acids
and ammonium
hydroxide

K E Y F A C T S

OVERVIEW

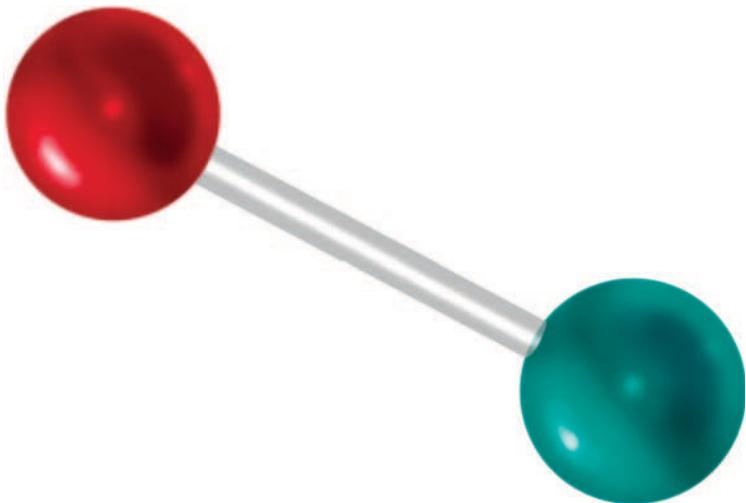
Copper(II) oxide (KOPP-er two OK-side) occurs in nature in the minerals tenorite, melaconite, and paramelaconite. In pure form, it is a black to brownish powder or crystalline material. Like copper(I) oxide, copper(II) oxide is a semiconductor, a material that conducts an electric current, although not nearly as well as conductors such as gold, silver, and aluminum.

HOW IT IS MADE

Copper(II) oxide forms naturally in the Earth as a result of the weathering of copper sulfides (Cu_2S and CuS). It is prepared synthetically by heating copper metal in air to about 800°C (1,500°F) or, more commonly, by heating copper(II) carbonate (CuCO_3) or copper(II) nitrate [$\text{Cu}(\text{NO}_3)_2$] to red heat.

COMMON USES AND POTENTIAL HAZARDS

Throughout recorded history, copper(II) oxide has been used as a pigment to color ceramics, enamels, porcelain



Copper(II) oxide. Red atom is oxygen and turquoise atom is copper. PUBLISHERS RESOURCE GROUP

glazes, and artificial gems, applications that continue to the present day. The oxide adds a bluish to greenish tint to such materials. Copper(II) oxide also finds use as an insecticide and fumigant. It is used primarily in the treatment of potato plants and as an antifouling agent on boat hulls. An antifouling agent is a material that prevents the formation of barnacles and other organisms on the bottom of a boat. When such organisms grow on a boat's hull, they increase the friction produced when the boat rides through the water, thus reducing its speed. The compound is also used as a wood preservative, to protect fence posts, pilings, decking, roofing, shingles, sea walls, and other freshwater and marine structures from insects and fungi.

Other uses to which copper(II) oxide is put including the following:

- In the preparation of superconducting materials, materials that have essentially no resistance to the flow of an electric current;
- In the manufacture of batteries and electrodes;
- As a welding flux for use with bronze objects and materials;
- For polishing of optical glass, glass used in telescopes, microscopes, and similar instruments;
- In the preparation of phosphors, materials that glow in the dark after being exposed to light;

Interesting Facts

- The ancient Greeks used a mixture of copper(II) oxide and copper(II) sulfate to treat wounds.
- U.S. pennies sometimes develop a black coating that is caused by copper(II) oxide. The coating can be removed by cleaning the penny in a solution of vinegar, lemon juice, and salt. Coca Cola™ can also be used because it contains phosphoric acid (H_3PO_4), which dissolves copper(II) oxide.
- Some fingerprinting powders contain copper(II) oxide.

- For the removal of sulfur and sulfur compounds from petroleum;
- In the manufacture of rayon; and
- As a catalyst in many industrial and commercial chemical reactions.

Inhaling copper(II) oxide fumes may result in a condition known as metal fume fever, with irritation of the throat, coughing, shortness of breath, nausea, and fever. Excessive exposure may result in chronic lung disease. Ingesting large amounts of copper(II) oxide may cause vomiting, diarrhea, nausea, excessive salivation, and intense abdominal pain. The compound is also an eye irritant.

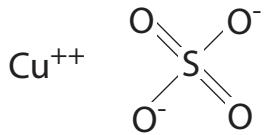
Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

FLUX A material that lowers the melting point of another substance or mixture of substances or that is used in cleaning a metal.

FOR FURTHER INFORMATION

- “Copper Oxide 99.99+%.” Louisiana State University.
http://www.camd.lsu.edu/msds/c/copperII_oxide.htm#Synonyms
(accessed on October 5, 2005).
- “Cupric Oxide.” J. T. Baker.
<http://www.jtbaker.com/msds/englishhtml/c5885.htm>
(accessed on October 5, 2005).
- Richardson, H. W., ed. *Handbook of Copper Compounds and Applications*. New York: Marcel Dekker, 1997.



OTHER NAMES:

Cupric sulfate; blue vitriol; blue stone; blue copperas

FORMULA:

CuSO_4

ELEMENTS:

Copper, sulfur, oxygen

COMPOUND TYPE:

Inorganic salt

STATE:

Solid

MOLECULAR WEIGHT:

159.61 g/mol

MELTING POINT:

Not applicable; decomposes above 560°C ($1,040^\circ\text{F}$)

BOILING POINT:

Not applicable

SOLUBILITY:

Very soluble in water; moderately soluble in methyl alcohol; slightly soluble in ethyl alcohol

**K
E
Y
F
A
C
T
S**

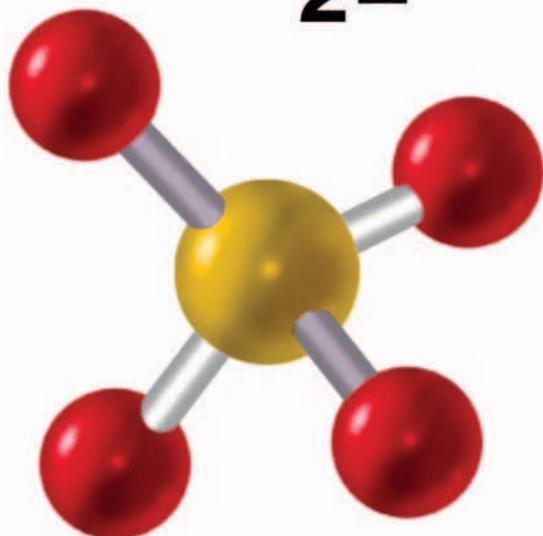
OVERVIEW

Copper(II) sulfate (KOPP-er two SUL-fate) is a white crystalline powder in its anhydrous state, although it occurs most commonly as the pentahydrate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$, which is a blue granular crystalline solid. When the pentahydrate is heated, it loses its water of hydration and returns to its white, powdery anhydrous state. The transition between its white anhydrous form and its blue hydrated form is sometimes used in devices and toys that indicate the presence or absence of moisture in the air.

Copper(II) sulfate is probably the most important and most widely used of all copper compounds. It is used in agriculture, the paints and covering industry, in electrical applications, in the production of other chemicals, and in other industrial and commercial applications.

HOW IT IS MADE

Copper(II) sulfate occurs in nature as the mineral calcanthite. Calcanthite is somewhat rare since copper(II)

2+**2-**

Copper(II) sulfate pentahydrate. Red atoms are oxygen; yellow atom is sulfur; and turquoise atom is copper. Gray sticks are double bonds.

PUBLISHERS RESOURCE GROUP

sulfate dissolves readily in water, and therefore the mineral can leach out of rock by groundwater. The compound is made in the laboratory and commercially by treating copper metal or one of the copper oxides (CuO or Cu_2O) with sulfuric acid (H_2SO_4).

COMMON USES AND POTENTIAL HAZARDS

Copper(II) sulfate is used extensively in agriculture as a soil additive, a fumigant for trees, a feed additive to prevent mineral deficiencies, and as a wood preservative. It is a component, along with calcium oxide (CaO ; quicklime), of one of the oldest and most popular plant fumigants, a material called Bordeaux mixture. Many of copper(II) sulfate's uses depend on its highly toxic character. It is used to prevent the growth of mold and mildew; to sanitize surfaces in hospitals; to kill algae in swimming pools, reservoirs, and other bodies of water; as an herbicide to kill weeds; and as a germicide, to prevent the growth of unwanted seeds. Although very effective in such applications, copper(II) is often not the first choice because it may have toxic effects on other plants and animals in the environment and on humans.

Interesting Facts

- Copper(II) sulfate gets the name blue vitriol from the color of the pentahydrate (blue) plus the fact that it is made from sulfuric acid, which is also called oil of vitriol.

Some additional uses of copper(II) sulfate include the following:

- As a mordant in the textile industry;
- In the production of blue pigments for paints, varnishes, glazes, dyes, inks, and other coloring materials;
- In electroplating, where it supplies the copper with which other metals are plated;
- In making lithographic prints and engravings;
- In the preservation and tanning of animal hides;
- As a catalyst in the refining of petroleum;
- As a nutritional supplement to treat copper deficiency disorders;
- In the manufacture of fireworks, where it imparts a blue color to the display; and
- In the production of water-resistant adhesives for wood.

Copper(II) sulfate is known to be toxic to humans and other animals if ingested. It may cause burning in the mouth and stomach, nausea, vomiting, abdominal pain, diarrhea, and a metallic taste in the mouth. It can cause gastrointestinal bleeding and, if not eliminated from the body, can result in kidney and liver damage, depression of the nervous system, paralysis, coma, and death from shock or kidney failure. Prolonged exposure to copper(II) sulfate dust can cause discoloration of the skin, damage to the blood and liver, and irritation of the eyes and nose. Copper(II) sulfate poses some risk to the environment, both because of its toxic nature and because it does not break down readily.

Words to Know

ANHYDROUS Form of a compound that lacks water.

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

FUMIGANT A pesticide that is applied as a gas.

MORDANT A substance used in dyeing and printing that reacts chemically with both a dye and the material being dyed to help hold the dye permanently to the material.

FOR FURTHER INFORMATION

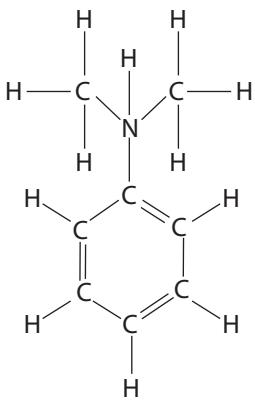
“Cupric Sulfate.” New Jersey Department of Health and Senior Services.
<http://www.state.nj.us/health/eoh/rtkweb/o549.pdf>
(accessed on October 5, 2005).

“Cupric Sulfate Anhydrous.” J. T. Baker.
<http://www.jtbaker.com/msds/englishhtml/c5920.htm>
(accessed on October 5, 2005).

Ware, George W. *The Pesticide Book*. Batavia, Ill.: Mesiter, 1999.

Young, Jay A. “Copper (II) Sulfate Pentahydrate.” *Journal of Chemical Education* (February 2002): 158.

See Also Copper(II) Oxide



OTHER NAMES:
Isopropylbenzene; 1-(methylethyl) benzene
FORMULA:
$C_6H_5CH(CH_3)_2$
ELEMENTS:
Carbon, hydrogen
COMPOUND TYPE:
Aromatic hydrocarbon (organic)
STATE:
Liquid
MOLECULAR WEIGHT:
120.19 g/mol
MELTING POINT:
-96.02°C (-140.8°F)
BOILING POINT:
152.41°C (306.34°F)
SOLUBILITY:
Insoluble in water; miscible with ethyl alcohol, benzene, ether, acetone, and most organic solvents

KEY FACTS

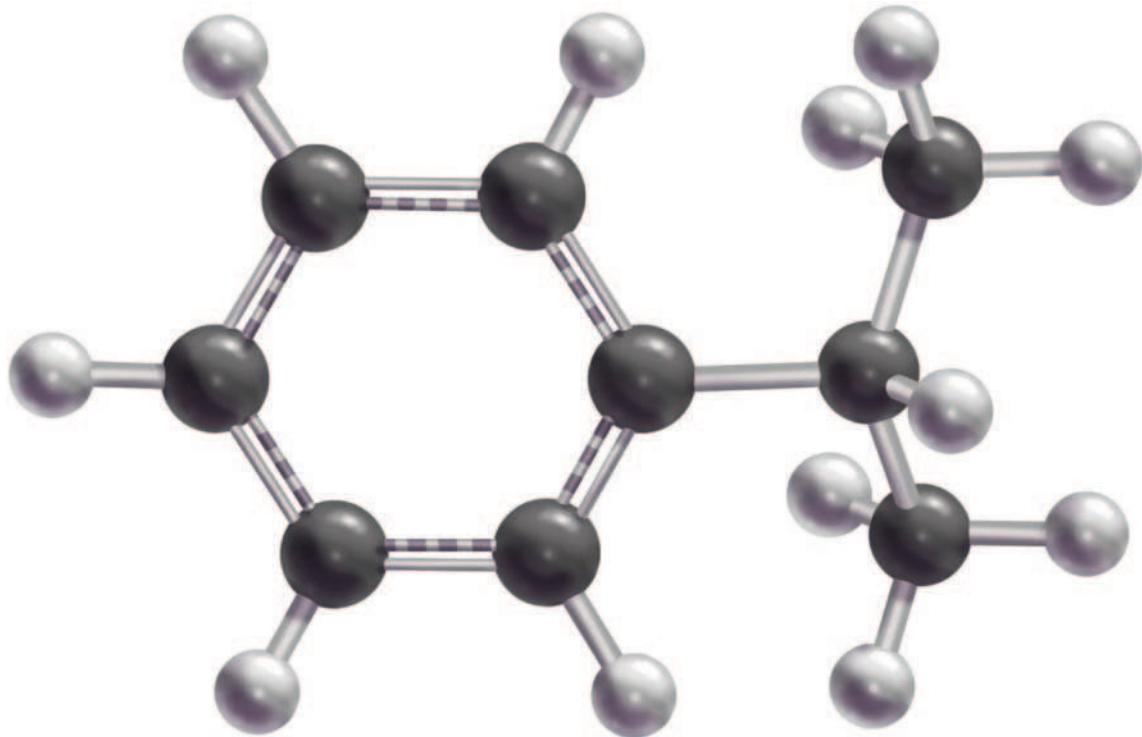
Cumene

OVERVIEW

Cumene (KYOO-meen) is a colorless, flammable liquid with a penetrating gasoline-like odor. Chemically, it is classified as an aromatic hydrocarbon. Aromatic hydrocarbons are compounds of carbon and hydrogen with a molecular structure based on that of benzene. Although cumene is probably not well known to the average person, it is a very important industrial chemical. In 2004, some 3.74 million metric tons (4.12 million short tons) of cumene were produced in the United States, making it the twentieth most important chemical made in that year by weight. Cumene is used primarily as a raw material in the synthesis of other organic compounds, such as phenol, acetone, acetophenone, and methyl styrene, and as a thinner for paints and lacquers.

HOW IT IS MADE

Cumene is made in a straight-forward chemical reaction between benzene (C_6H_6) and propene (propylene; $CH_3CH=CH_2$).



Cumene. Black atoms are carbon; white atoms are hydrogen. Bonds in the benzene ring are represented by the double striped sticks.

White sticks show single bonds. PUBLISHERS RESOURCE GROUP

It can also be extracted from petroleum or from coal tar, the thick black material left after soft coal is converted to coke.

COMMON USES AND POTENTIAL HAZARDS

Small amounts of cumene are used as thinners for paints, lacquers, and enamels, and as solvents in paints and other types of coatings. By far the greatest amount of cumene, however, is used as a raw material in the manufacture of phenol, acetone, and methyl styrene. These compounds, in turn, have a great many chemical and industrial uses. Some of the most important uses are the production of plastics, such as polystyrene, phenol-formaldehyde resins, and polycarbonates.

Cumene is a skin, eye, and respiratory system irritant. If inhaled, it can cause coughing, dizziness, drowsiness,

Interesting Facts

- Traditionally, the synthesis of cumene has involved the use of phosphoric acid (H_3PO_4) as a catalyst. That process results in the release of harmful by-products into the environment, and a new process that uses

zeolites as a catalyst has now become more popular. Zeolites are a naturally occurring, earthy material that can be as effective as phosphoric acid in promoting the synthesis of cumene.

sore throat, headache, and loss of muscular coordination. In large doses, it has a narcotic effect, resulting in drowsiness and insensitivity to pain and other stimuli, and may lead to unconsciousness. There is no evidence that cumene is carcinogenic or that it causes hereditary damage to a person or to offspring of individuals exposed to the compound.

Words to Know

AROMATIC HYDROCARBON A compound of carbon and hydrogen with a molecular structure based on that of benzene.

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

MISCIBLE Able to be mixed; especially applies to the mixing of one liquid with another.

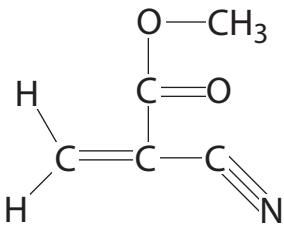
SOLVENT A substance that is able to dissolve one or more other substances.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

FOR FURTHER INFORMATION

- “Cumene.” International Programme on Chemical Safety.
[http://www.inchem.org/documents/cicads/cicads/cicad18.htm#
PartNumber:1](http://www.inchem.org/documents/cicads/cicads/cicad18.htm#PartNumber:1) (accessed on December 22, 2005).
- “Cumene.” Technology Transfer Network, Air Toxics Website
<http://www.epa.gov/ttn/atw/hlthef/cumene.html> (accessed on December 22, 2005).
- “Cumene (1 methylethylbenzene).” Australian Government. Department of the Environment and Heritage.
http://www.npi.gov.au/database/substance_info/profiles/28.html (accessed on December 22, 2005).

See Also Acetone; Phenol; Styrene



OTHER NAMES:

See Overview; data below are for the methyl ester.

FORMULA:

CH₂=C(CN)COOCH₃

ELEMENTS:

Carbon, hydrogen, oxygen, nitrogen

COMPOUND TYPE:

Ester (organic)

STATE:

Liquid

MOLECULAR WEIGHT:

111.10 g/mol

MELTING POINT:

-22°C (-7.6°F)

BOILING POINT:

47°C (120°F)

SOLUBILITY:

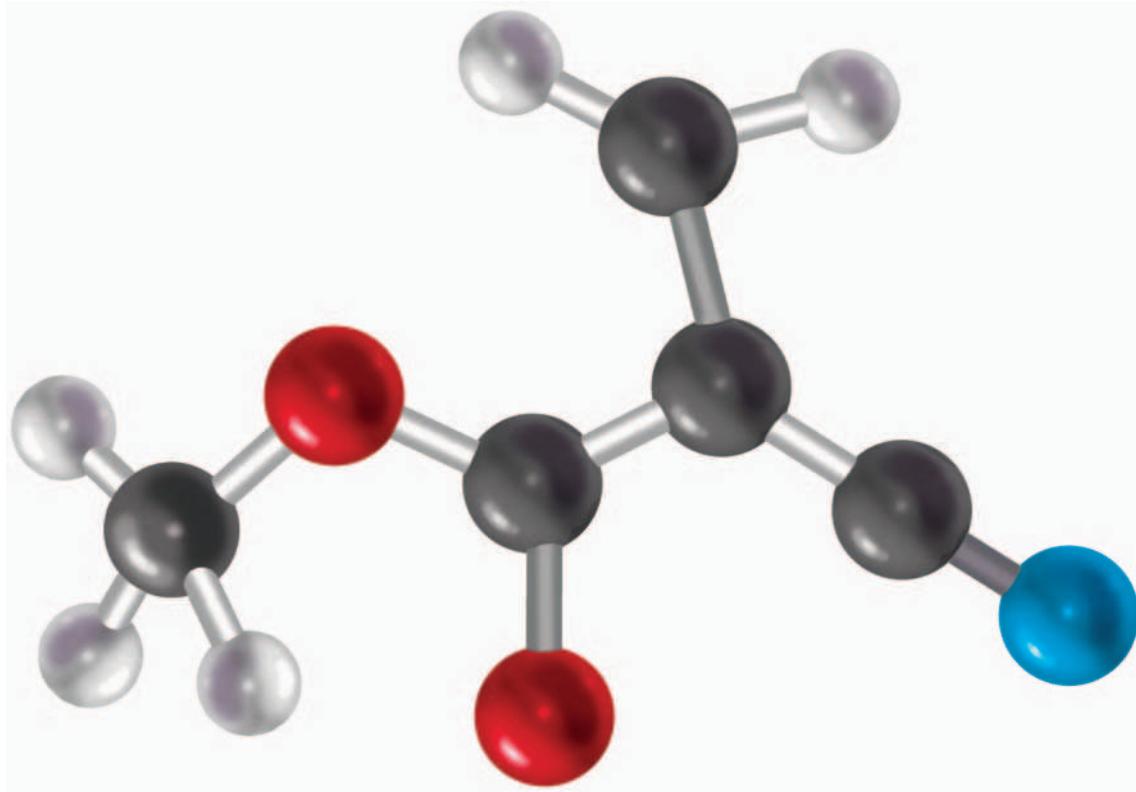
Reacts with water; soluble in acetone, toluene, methylene chloride, and some other organic solvents

K E Y F A C T S

OVERVIEW

Cyanoacrylate (sy-e-AN-oh-ACL-ri-late) is a general term used to describe a family of esters derived in the reaction between cyanoacrylic acid and various alcohols, such as methanol (methyl-2-cyanoacrylate) and ethanol (ethyl-2-cyanoacrylate). Cyanoacrylates are all esters of cyanoacrylic acid, CH₂=C(CN)COOH, and include methyl-2-cyanoacrylate, ethyl-2-cyanoacrylate, and 2-octyl cyanoacrylate. The material is the primary ingredient in Super Glue™, Krazy Glue™, Duro™, Quick Gel™, and a number of other products used to glue two surfaces together with a very strong bond in a matter of seconds. The bonding activity of the cyanoacrylates occurs when the esters react with water to form polymers that bond to materials with which they come into contact. The polymerization reaction occurs very quickly, resulting in strong bonds that form almost instantaneously.

Cyanoacrylate was discovered in 1942 by American researchers Harry Coover (1919-) and Fred Joiner, employees of the Eastman Kodak Laboratories, in Rochester, New York.



Cyanoacrylate. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atom is nitrogen. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

Coover and Joiner were attempting to synthesize a clear plastic material that could be used for gun sights. Instead, they obtained a liquid with a tendency to stick tightly to anything with which it came into contact. They decided that the product—then known as Eastman 910—was more of a problem than an asset and set it aside for six years before deciding to explore its adhesive properties.

At that point, Eastman Kodak began studying the possibilities of using cyanoacrylate for military applications, specifically for sealing wounds. It applied to the U.S. Food and Drug Administration (FDA) for permission to market the product for this use, an application that was approved in 1964. Shortly thereafter, military physicians tested cyanoacrylate glues in Vietnam and found them highly effective in sealing the wounds of soldiers. In a spray form, the glue stopped bleeding from chest wounds long enough to allow

wounded soldiers to reach a hospital. Although the use of cyanoacrylates by the military proved to be a success, the FDA declined to approve the product for civilian use. The product had a tendency to irritate the skin when it reacted with moisture on the skin, releasing formaldehyde and cyanoacetate, both toxic substances. Eventually, research showed that the octyl ester, 2-octyl-cyanoacrylate, proved to form stronger bonds, was more flexible, and was less irritating than the methyl and ethyl esters used in the original Eastman 610 formulation. With this change, the FDA approved the use of 2-octyl-cyanoacrylate for use in closing wounds and incisions. In 2001, the FDA extended this approval to include use of the compound as an antibacterial agent.

HOW IT IS MADE

Cyanoacrylates are made by reacting cyanoacrylic acid with an alcohol, such as methanol, ethanol, or 2-octanol, to form the corresponding ester.

COMMON USES AND POTENTIAL HAZARDS

Cyanoacrylates are still used in a variety of medical applications to close wounds and to protect wounds from infection. The 2-octyl ester is used for this purpose because it has fewer hazardous side effects than do the methyl and ethyl esters. Cyanoacrylate is preferred to stitches for closing wounds because it is less likely to become infected and tends to leave fewer scars. For this reason, the compound is often used for cuts on the face, although it cannot be used in the nose or mouth. Consumers can buy cyanoacrylate products such as BandAid brand Liquid BandageTM for home use.

Cyanoacrylates are sometimes used in dental procedures. Their ability to bond tightly and quickly, to seal wounds, and to prevent the development of infections make them useful for a number of surgical procedures involving teeth and gums. The primary concern is the selection of esters (such as the 2-octyl ester) that will not cause irritation or infections of the areas being treated.

Cyanoacrylates are also used widely in a number of industrial and household adhesives, perhaps the best known of which is Super GlueTM. Manufacturers use cyanoacrylates in the assembly of many products, including trophies, golf

Interesting Facts

- The commercial success of cyanoacrylates as glues was assured in 1959 when Harry Coover appeared on the television show *I've Got a Secret* and used the product to lift host Gary Moore off the floor with a single drop of Eastman Kodak's Super Glue™.
- The polymerization of cyanoacrylates with water

begins as soon as the glue is spread on a material, and a strong bond is formed almost immediately. However, polymerization continues for a number of hours and is not complete for 24 to 48 hours, at which point the bonding is considerably stronger than in the first few minutes of application.

clubs and other sports equipment, tools, digital watches, optical lenses, electronic components, lampshades, plastics, scientific instruments, loudspeakers, shoes, jewelry, and videocassettes. Automobile makers use the product to attach trim to cars and trucks. The cyanoacrylates come in a great variety of formulations that make them suitable for bonding with special types of materials (such as ceramics, metals, wood, plastics, and glasses) and under a variety of specialized conditions (such as high or low temperature or very dry or damp conditions).

Cyanoacrylates are also used by forensic scientists to collect fingerprints at crime scenes. The object to be tested is suspended inside a container with at least one transparent side. A few drops of Super Glue or similar cyanoacrylate product is added to the container and the container is sealed and heated to about 100°C (212°F). That heat causes vaporization and polymerization of the cyanoacrylate, resulting in the formation of distinctive white print patterns, a process that may take two hours or more. This cyanoacrylate fuming test has now become the procedure of choice for the detection of latent prints deposited on non-porous objects, such as glass, plastic, rubber, and leather.

Cyanoacrylates bond to skin as quickly and tightly as they do to other materials, so care is required in their use.

Words to Know

ESTER An organic compound formed in the reaction between an organic acid and an alcohol.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

POLYMERIZATION The process of creating a polymer; a reaction that causes a polymer to form.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

They are skin and eye irritants, and so should always be used in an area with good ventilation. Exposure to high levels of cyanoacrylate gas can cause severe respiratory problems, and in the most extreme cases can result in death.

FOR FURTHER INFORMATION

“Cyanoacrylates.” Impact Adhesives.

<http://www.impactadhesives.com/pages/cyano.html> (accessed on October 6, 2005).

Davidson, Elaine. “Cyanoacrylates.”

<http://www.repairantiques.com/cyanoacrylates.html> (accessed on October 6, 2005).

Fernandez, Tania, and Val Bliskovsky. “Cyanoacrylate Technology: Stay Glued.” Pharmabiz.com.

<http://www.pharmabiz.com/article/detnews.asp?articleid=13609§ionid=46> (accessed on October 6, 2005).

“Methyl α -Cyanoacrylate (MCA); Ethyl α -Cyanoacrylate (ECA).” Occupational Safety and Health Administration.

<http://www.osha.slc.gov/dts/sltc/methods/organic/orgo55/orgo55.html> (accessed on October 6, 2005).

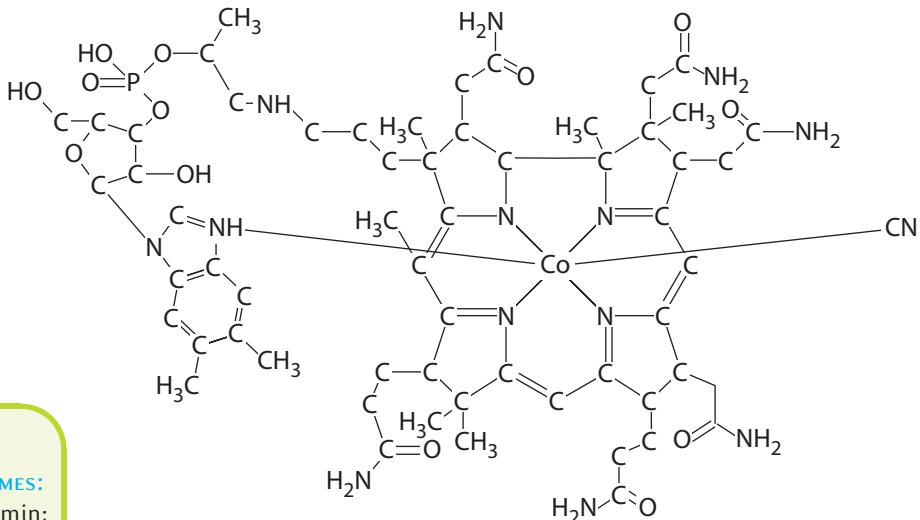
“Methyl Cyanoacrylate and Ethyl Cyanoacrylate.” International Programme on Chemical Safety.

<http://www.inchem.org/documents/cicads/cicads/cicad36.htm#9.0> (accessed on October 6, 2005).

Schwade, Nathan D. “Wound Adhesives: α -Octyl Cyanoacrylate.” eMedicine: Instant Access to Medicine.

<http://www.emedicine.com/ent/topic375.htm> (accessed on October 6, 2005).

See Also Formaldehyde



OTHER NAMES:

Cobalamin;
vitamin B₁₂

FORMULA:

C₆₃H₈₈CoN₁₄O₁₄P

ELEMENTS:

Carbon, hydrogen,
cobalt, nitrogen,
oxygen, phosphorus

COMPOUND TYPE:

Heterocyclic ring
(organic)

STATE:
Solid

MOLECULAR WEIGHT:
1355.36 g/mol

MELTING POINT:
Undetermined;
greater than 300°C
(575°F)

BOILING POINT:
Not applicable

SOLUBILITY:
Soluble in water in
alcohol; insoluble in
acetone, ether, and
chloroform

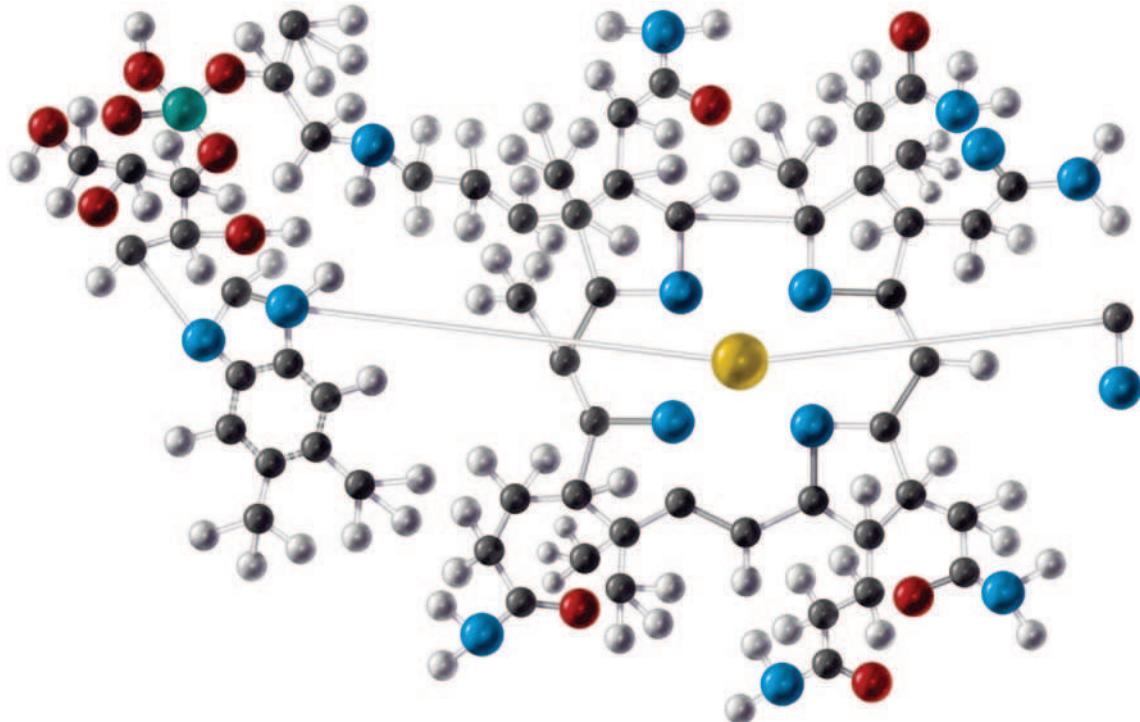
K E Y F A C T S

OVERVIEW

Cyanocobalamin (sye-AN-oh-koh-BAL-uh-min) is more commonly known as vitamin B₁₂. At least three active forms of the vitamin are known. They include hydroxocobalamin and nitrocobalamin, in addition to cyanocobalamin, all with slightly different molecular structures. Cyanocobalamin occurs as dark red crystals or a red powder that is odorless and tasteless. When heated, the compound darkens above 200°C (392°F), but does not melt when heated further.

Cyanocobalamin is required for the synthesis of DNA in cells and for the proper functioning of red blood cells and nerves. A deficiency of vitamin B₁₂ in the diet results in a condition known as pernicious anemia, a condition in which red blood cells are smaller in size, number, and hemoglobin content than normal red blood cells. These defects result in general weakness, nervous disorders, gastrointestinal disturbances, and, eventually and inevitably, death.

CYANOCOBALAMIN



Cyanocobalamin. Blue atoms are nitrogen; yellow atom is cobalt; green atom is phosphorous; black atoms are carbon, white atoms are hydrogen; red atoms are oxygen. Gray sticks are double bonds; striped sticks show benzene ring.

PUBLISHERS RESOURCE GROUP

The earliest research on pernicious anemia dates to the 1920s when American physician George Hoyt Whipple (1878-1976) found that anemia in dogs could be cured by feeding them beef liver. Inspired by this work, American physicians William Murphy (1892-1987) and George Minot (1885-1950) discovered that feeding pernicious anemia patients beef liver cured the disease. These discoveries led to the suspicion that one of the B vitamins present in liver was responsible for the cure of pernicious anemia. Scientists even invented a name for the missing vitamin, vitamin B₁₂, long before it was actually discovered. The problem, unknown to researchers at the time, was that vitamin B₁₂ was such a complex molecule that its isolation and identification was a significant challenge.

In 1948, American chemist Karl Folkers (1906-1997) and his research team discovered that they could measure

Interesting Facts

- Cyanocobalamin gets its name because it contains the metal cobalt.
- For a period of time, the only cure available for pernicious anemia was for patients to eat about a

pound of liver every day. Since many patients disliked that diet, they were very happy when synthetic vitamin B₁₂ supplements became available.

amounts of the vitamin by measuring the growth rate of bacteria that grew on vitamin B₁₂. Soon the team learned how to purify the vitamin, which formed small red crystals that many scientists consider quite beautiful. It still took many years for scientists to determine exactly how the enormous molecule was structured. In 1956, however, the British chemist Dorothy Hodgkin (1910-1994) determined its chemical structure, allowing her American colleague Robert Burns Woodward (1917-1979) to actually synthesize the compound, an accomplishment of enormous complexity that took fifteen years to complete.

HOW IT IS MADE

Vitamin B₁₂ is made naturally by bacteria that live in the intestines of all animals, including humans, as well as in soil. It binds to protein in food. Plants do not synthesize vitamin B₁₂. Manufacturers who make vitamin B₁₂ supplements use bacteria to grow the vitamin by a process similar to that which occurs naturally. Good food sources of vitamin B₁₂ include animal foods, such as fish, meat, poultry, eggs, milk, cheese, and yogurt; as well as fortified cereals. People who do not eat animal products should be sure to select foods fortified with artificially produced vitamin B₁₂, because the synthetic vitamin is produced by a natural process that does not involve the destruction of animals or the consumption of animal products.

Words to Know

DNA Deoxyribonucleic acid, a material with a cell that carries its genetic information and is capable of reproducing itself.

HETEROCYCLIC RING A compound whose molecules contain at least one ring in

which some element other than carbon is also present.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

COMMON USES AND POTENTIAL HAZARDS

Cyanocobalamin is essential to human health. The body cannot make properly functioning red blood cells, white blood cells, or platelets without it. It is also essential for the production of DNA and nerve cells. People whose diet is deficient in cyanocobalamin are likely to develop pernicious anemia. The symptoms of this disorder include fatigue, weakness, numbness, blurred vision, poor memory, confusion, hallucinations, personality changes, a smooth shiny tongue, stiffness, diarrhea, poor appetite, slow growth in children, and reduced sensitivity to pain and pressure.

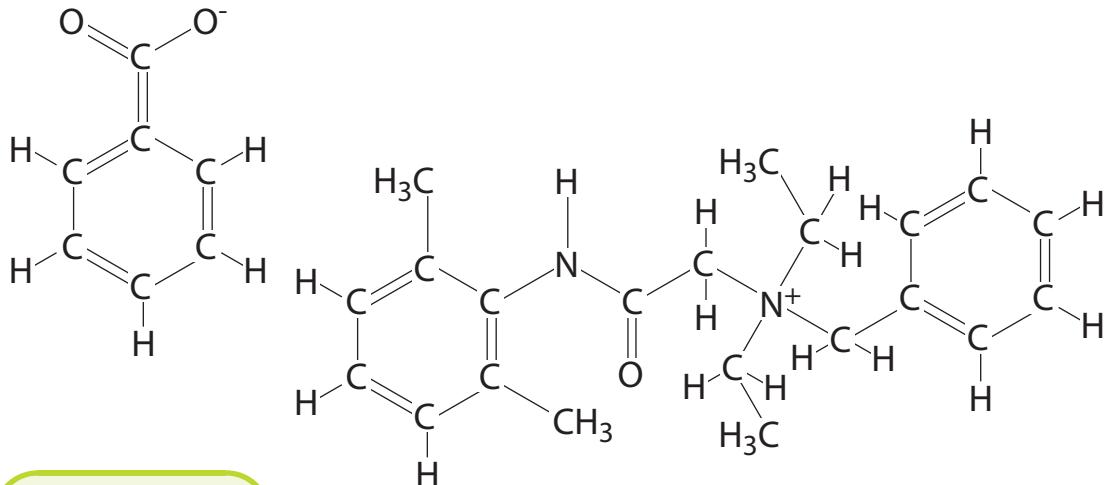
FOR FURTHER INFORMATION

Brody, Tom. *Nutritional Biochemistry*. San Diego: Academic Press, 1998.

“Dietary Supplement Fact Sheet: Vitamin B₁₂,” National Institutes of Health, Office of Dietary Supplements.
<http://ods.od.nih.gov/factsheets/vitaminb12.asp> (accessed on October 6, 2005).

“Nascobal Gel.” Nastech Pharmaceutical Company.
<http://www.tzamal medical.co.il/tzamal pharma products nascobal.htm> (accessed on October 6, 2005).

Opt, Robert C., and David J. Brown. "Vitamin B₁₂ Deficiency." *American Academy of Family Physicians*. March 1, 2003. Also available online at <http://www.aafp.org/afp/20030301/979.html> (accessed on October 6, 2005).



OTHER NAMES:

Benzinemethanamium;
other names are
possible

FORMULA:

$C_{28}H_{34}N_2O_3$

ELEMENTS:

Carbon, hydrogen,
nitrogen, oxygen

COMPOUND TYPE:

Organic acid

STATE:

Solid

MOLECULAR WEIGHT:

446.58 g/mol

MELTING POINT:

166°C to 170°C (331°F
to 338°F)

BOILING POINT:

Not applicable

SOLUBILITY:

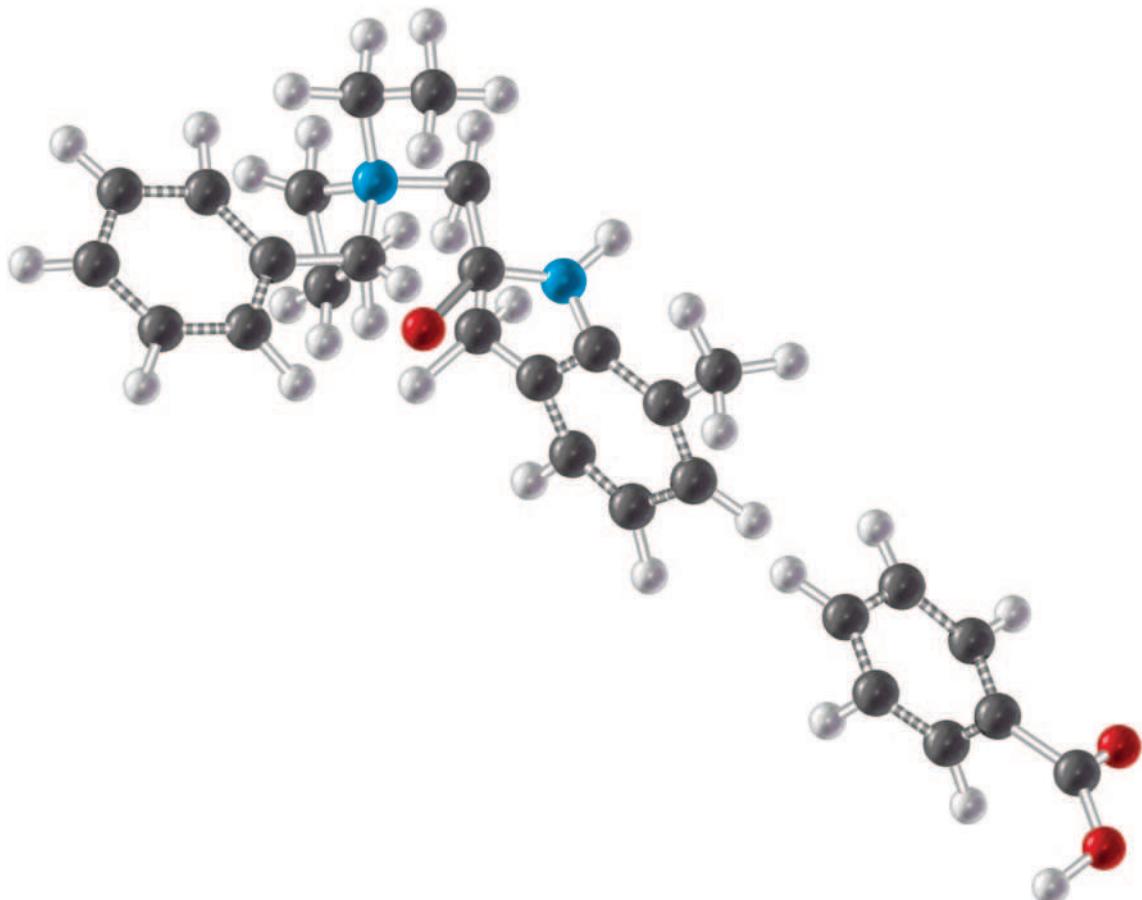
Soluble in water and
alcohol; insoluble in
ether

K E Y F A C T S

OVERVIEW

Denatonium benzoate (de-an-TOE-nee-um BEN-zoh-ate) is generally regarded as having the most bitter taste of any compound known to science. It is sold under the trade name of Bitrex®. Although denatonium benzoate has a powerful taste, it is colorless and odorless. The taste is so strong, however, that most people cannot tolerate a concentration of more than 30 parts per million of denatonium benzoate. Solutions of denatonium benzoate in alcohol or water are very stable and retain their bitter taste for many years. Exposure to light does not lessen the compound's bitter taste.

Denatonium benzoate compound was discovered in 1958 by a scientist named W. Barnes, who was working for the chemical firm of T. & H. Smith, in Edinburgh, Scotland. Barnes was interested in developing a new anesthetic, more powerful than those already available to physicians. He decided to focus his research on lidocaine, a very popular anesthetic, and compounds chemically related to it. In one line of his experiments, Barnes added a single benzoyl group



Denatonium Benzoate. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atoms are nitrogen. Gray sticks indicate double bonds; striped sticks show benzene rings.

PUBLISHERS RESOURCE
GROUP

(benzoic acid with a hydrogen removed: $\text{C}_6\text{H}_5\text{COO}^-$) to a nitrogen atom in lidocaine. The resulting compound was denatonium benzoate. Although the compound had little effectiveness as an anesthetic, Barnes noted that it had a peculiar odor and taste. The Smith company decided to exploit this unusual property of denatonium benzoate, and obtained a patent for it under the name of Bitrex®. Today, the primary manufacturer of Britex® in the world is the Macfarlan Smith corporation of Edinburgh.

HOW IT IS MADE

The process by which denatonium benzoate is made is a proprietary secret of the Macfarlan Smith corporation. A proprietary secret is a method of making a product for which

Interesting Facts

- The claim that denatonium benzoate is the most bitter tasting chemical known is based not on scientific tests but on human taste tests alone. No automated test exists for determining the bitterness of a substance.
- Oregon was the first state in the United States to require the addition of Bitrex® to antifreeze and car windshield washer fluid. The requirement was instituted to prevent people from drinking such products accidentally or, in the case of alcoholics, intentionally.

a company holds a patent and the details of which it does not disclose to the general public.

COMMON USES AND POTENTIAL HAZARDS

One of the first and most important uses of denatonium benzoate was as an additive to methanol (methyl alcohol; wood alcohol). Although ethanol (ethyl alcohol; grain alcohol) has some harmful effects on humans, especially if taken in excess, it is relatively safe to drink in beer, wine, and other alcoholic drinks. By contrast, methanol is highly toxic. Any one who accidentally or intentionally consumes methanol is likely to experience serious health effects, including death. By adding a small amount of denatonium benzoate to methanol, consumers are discouraged—and usually prevented—from drinking the substance.

Denatonium benzoate has many other applications. For example, it can be used in a dilute solution to brush on the fingernails of people who are compulsive fingernail-biters. Some parents use a similar solution on the thumbs of children who suck their thumbs more than they should. Denatonium benzoate is also used as an animal repellent. Products containing denatonium benzoate can be sprayed on trees, brushes, crops, and other material to prevent deer from grazing on those products. One of the product's first applications was as a treatment on pig's tails to prevent pigs from biting

each other. The coatings on electric cables are sometimes impregnated with a denatonium benzoate solution to discourage rats from chewing on them.

Some of the other applications in which denatonium benzoate has been used include the following:

- In liquid laundry detergents;
- In fabric conditioners;
- In toilet cleaners;
- In disinfectants;
- In household antiseptics;
- In kitchen, bathroom, and floor cleaners;
- In paint products and paint brush cleaners;
- In personal care products, including bath foam, soaps, perfume and after shave lotions, nail polish remover, shampoo, and shower gel;
- In pesticides, such as insecticides, rodenticides, slug bait, and ant bait;
- In herbicides; and
- In a wide variety of automotive care products, such as antifreezes, coolants, and car cleaning materials.

In all of these cases, the purpose of adding denatonium benzoate is to change the taste of the product just enough to prevent someone, especially children, from eating a substance that could cause them harm.

Despite its bitter taste, denatonium benzoate appears to pose little or no hazard to human health. Exposure to the pure compound may cause respiratory discomfort, but only people working with the substance directly are likely to encounter this problem.

FOR FURTHER INFORMATION

“About Bitrex®.” Market Actives.

<http://www.marketactives.com/home.html> (accessed on October 7, 2005).

“Contains Bitrex.” Macfarlan Smith.

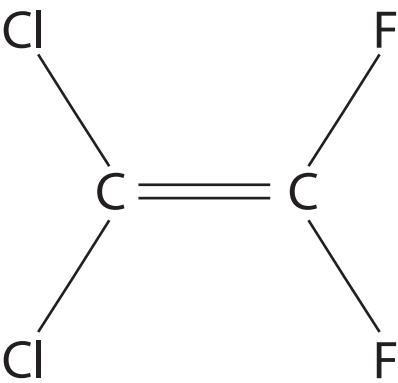
<http://www.bitrex.com/home.htm> (accessed on October 7, 2005).

"Denatonium Benzoate." C Tech Corporation.

<http://www.ctechcorporation.com/benzoate.htm> (accessed on October 7, 2005).

"The Most Bitter Substance." Center for the Science & Engineering of Materials.

http://www.csem.caltech.edu/material_of_month/bitrex.html (accessed on October 7, 2005).



OTHER NAMES:
Difluorodichloromethane; Freon 12®

FORMULA:
 CCl_2F_2

ELEMENTS:
Carbon, chlorine, fluorine

COMPOUND TYPE:
Halogenated hydrocarbon; alkyl halide (organic)

STATE:
Gas

MOLECULAR WEIGHT:
102.91 g/mol

MELTING POINT:
−158°C (−252°F)

BOILING POINT:
−29.8°C (−21.6°F)

SOLUBILITY:
Very slightly soluble in water; soluble in alcohol and ether

KEY FACTS

OVERVIEW

Dichlorodifluoromethane (DIE-klor-oh-DIE-floor-oh-METH-an)e belongs to a family of compounds called the chlorofluorocarbons (CFCs) that consist of carbon, chlorine, and fluorine atoms in varying arrangements. They can be considered as derivatives of alkanes, organic compounds consisting of carbon and hydrogen only, in which all of the hydrogen atoms have been replaced by chlorine and/or fluorine atoms. Chlorofluorocarbons are also referred to by their trade names, assigned to them by the DuPont corporation, which holds patents for their production. Dichlorodifluoromethane, for example, is also known as Freon12®. The numbers in each compound's Freon name specify the carbon atoms on which the chlorine and/or fluorine atoms are located. Freon12® is also known as refrigerant 12, propellant 12, and halon 122. The term halon refers to a related family of compounds consisting of carbon, chlorine, fluorine, bromine, and/or iodine.

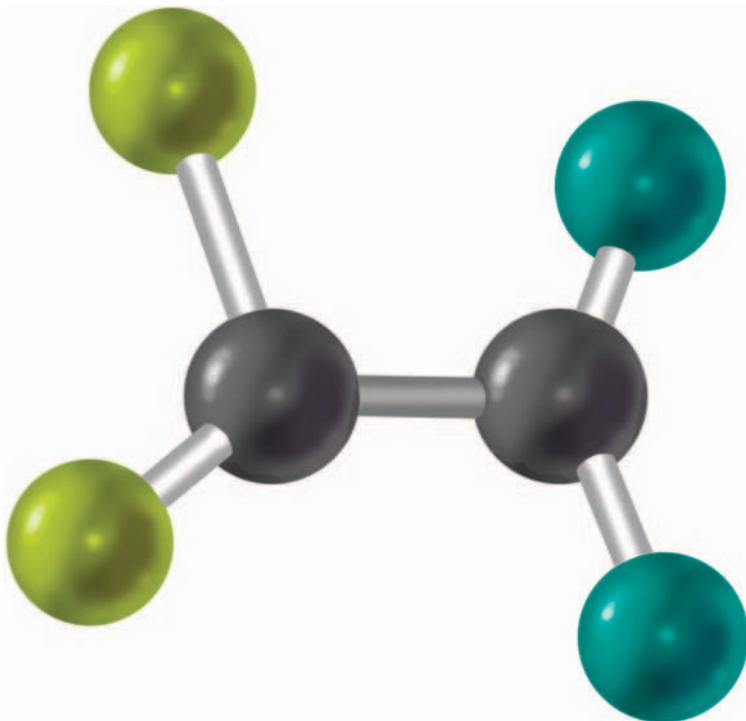
Dichlorodifluoromethane is the most widely used of all Freons®. It is a colorless, nonflammable gas that smells like ether in high concentrations.

DICHLORODIFLUOROMETHANE

Dichlorodifluoromethane.

Black atoms are carbon;
turquoise atoms are fluorine;
and green atoms are chlorine.
Gray stick shows a double bond.

PUBLISHERS RESOURCE GROUP



During the late 1800s and early 1900s, ammonia (NH_3) and sulfur dioxide (SO_2) were the gases most commonly used in refrigeration systems. These substances were toxic and flammable, however, and their use resulted in many deaths and injuries. Three U.S. companies, DuPont, General Motors, and Frigidaire, began searching for a nonflammable, non-toxic alternative to ammonia and sulfur dioxide in their refrigeration products. In the late 1920s, Thomas Midgley (1889–1944), a researcher at General Motors, discovered such a compound, dichlorodifluoromethane. It enabled cooling systems to operate more efficiently and more safely. DuPont, which had joined forces with General Motors in the search for a new coolant, began manufacturing Freon in 1930.

For several decades, Freon® was widely used in refrigeration and for a number of other industrial applications. In the 1970s, two scientists at the University of California, F. Sherwood Rowland (1927–) and Mario Molina (1943–) discovered that CFCs escaping from appliances on the Earth's surface drifted upward into the stratosphere, where they were exposed to intense ultraviolet radiation from the Sun.

That radiation caused the decomposition of CFCs into a number of products, one of which was chlorine. The chlorine produced in this breakdown, in turn, attacked ozone molecules (O_3) in the stratosphere and converted them into normal oxygen molecules (O_2).

Once recognized, this series of reactions became a matter of serious concern to scientists. The ozone layer absorbs ultraviolet radiation from the Sun, reducing the risk of skin cancer and other health problems produced by ultraviolet radiation. In 1987, twenty-seven nations had signed the Montreal Protocol on Substances that Deplete the Ozone Layer, which called for a 50-percent reduction in CFC production by the year 2000. An amendment to the protocol adopted in 1990 went even further, banning production of CFCs altogether in developed countries. Eventually, 148 countries were to sign the Montreal Protocol. By 1996, CFC production had come essentially to a halt except for a few developing nations and for use in certain specific applications, such as asthma inhalers. A decade later, scientists reported evidence that the ozone layer had begun to show signs of recovery.

Since the ban on CFC production has been adopted, manufacturers are replacing Freon® products with two related groups of compounds, the hydrochlorofluorocarbons (HCFCs) and hydrofluorocarbons (HFCs). These hydrogen-containing compounds are less energy efficient than the CFCs, but they are considered safer (although not entirely safe) for the environment.

HOW IT IS MADE

Dichlorodifluoromethane is made by reacting carbon tetrachloride (CCl_4) with hydrogen fluoride gas (H_2F_2) in the presence of a catalyst, usually antimony pentafluoride (SbF_5). In this reaction, two of the chlorines present in carbon tetrachloride are replaced by two fluorines from hydrogen fluoride, producing dichlorodifluoromethane.

COMMON USES AND POTENTIAL HAZARDS

Dichlorodifluoromethane and other Freons® were once used for many commercial applications: as refrigerants for air conditioners and refrigerators; as an aerosol propellant for hair sprays, insecticides, paints, adhesives, and cleaners;

Interesting Facts

- In 1930, Thomas Midgley demonstrated the safety of Freon® at a meeting of the American Chemical Society by breathing in the gas and then breathing it out to extinguish a candle flame.
- The production of dichlorodifluoromethane worldwide rose from about 150 million kilograms annually in 1960 to more than 750 million kilograms (1.65 billion pounds) annually in 1990. By 1996, because of the Montreal Protocol, it had dropped to less than 50 million kilograms (110 million pounds) annually.
- One chlorine atom produced in the stratosphere by the decomposition of a CFC molecule can destroy 100,000 molecules of ozone.

and as a foaming agent in the manufacture of shipping plastics. Its relative nontoxicity made these materials useful for food preservation and even for chilling cocktail glasses. Some other former uses of the Freons® include the following:

- As a rocket propellant;
- In solvents used in the manufacture of paints and varnishes;
- In the preparation of frozen sections of tissue;
- For the freezing of foods as a method of preservation;
- In water purification systems;
- As an agent for detecting leaks.

Companies favored Freons® for so many uses because they are both nontoxic and nonreactive in the troposphere. This property turned out to be a disadvantage also since it means that Freons® remain in the atmosphere for such long periods of time that they can eventually reach the stratosphere, where they pose a serious threat to the ozone layer.

Freons® may pose a health hazard if inhaled in large amounts. They displace oxygen in the air, causing dizziness, drowsiness, irregular heartbeat, cardiac arrest, unconscious-

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

STRATOSPHERE The layer of the Earth's atmosphere extending from about 15 kilometers (8 miles) to about 50

kilometers (30 miles) above the Earth's surface.

TROPOSPHERE The lowest layer of the Earth's atmosphere, reaching to about 8 miles (15 kilometers) above the Earth's surface.

ness, and death by suffocation. Sprayed on the skin, Freons® can cause frostbite, damage produced by exposure to freezing temperatures. The ignition of Freons® also poses hazards to human health because of the release of toxic gases such as hydrogen chloride (HCl), hydrogen fluoride (H₂F₂), and phosgene (COCl₂).

FOR FURTHER INFORMATION

Baker, Linda. "The Hole in the Sky (Ozone Layer)." *E* (November 2000): 34.

Cagin, Seth, and Philip Dray. *Between Earth and Sky: How CFCs Changed Our World and Endangered the Ozone Layer*. New York: Pantheon Books, 1993.

"Dichlorodifluoromethane." NIOSH Pocket Guide to Chemical Hazards.

<http://www.cdc.gov/niosh/npg/npgdo192.html> (accessed on October 7, 2005).

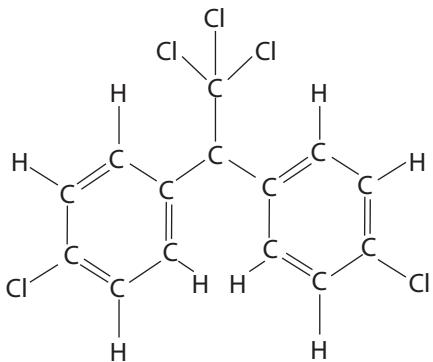
Elkins, James W. "Chlorofluorocarbons (CFCs)." National Oceanic and Atmospheric Administration (NOAA).

<http://www.cmdl.noaa.gov/noah/publictn/elkins/cfcs.html> (accessed on October 7, 2005).

"Gas Data." Air Liquide.

<http://www.airliquide.com/en/business/products/gases/gasdata/index.asp?GasID=22> (accessed on October 7, 2005).

See Also Ammonia; Hydrogen Chloride; Sulfur Dioxide



OTHER NAMES:
DDT; see Overview
for additional names

FORMULA:
 $(\text{ClC}_6\text{H}_4)_2\text{CHCCl}_3$

ELEMENTS:
Carbon, hydrogen,
chlorine

COMPOUND TYPE:
Halogenated
hydrocarbon
(organic)

STATE:
Solid

MOLECULAR WEIGHT:
354.49 g/mol

MELTING POINT:
108.5°C (227.3°F)

BOILING POINT:
260°C (500°F)

SOLUBILITY:
Insoluble in water;
slightly soluble in
ethyl alcohol; soluble
in ether, acetone, and
benzene

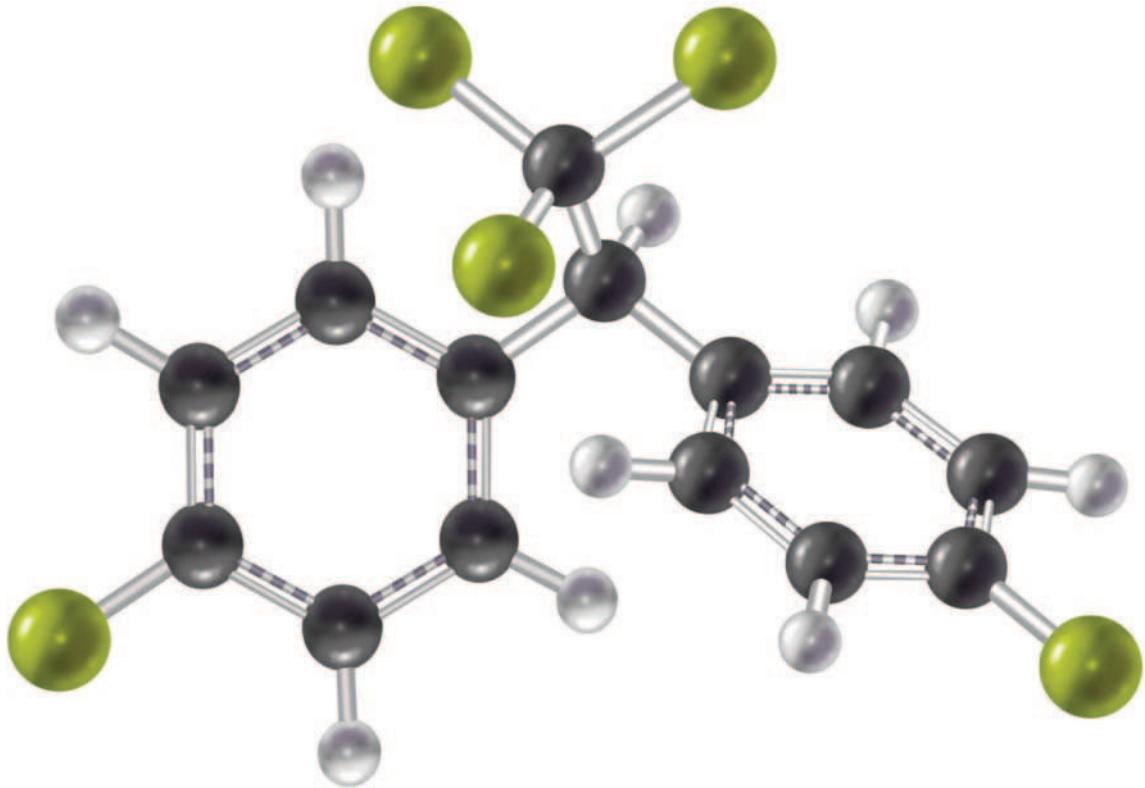
KEY FACTS

OVERVIEW

Dichlorodiphenyltrichloroethane (DI-klor-oh-DI-fee-nul-TRI-klor-oh-eth-an-e) is a colorless crystalline or white powdery material with a slight aromatic odor. It is far better known by its acronym, DDT. DDT was first synthesized in 1873 by German chemist Othmar Zeidler (1859–1911) as a project for his doctoral thesis. However, the compound was essentially ignored by other chemists and remained a laboratory curiosity for more than sixty years.

Then, in 1939, Swiss chemist Paul Hermann Müller (1899–1965) discovered that DDT is very effective in killing a wide variety of insects including the common housefly, the mosquito, the louse, and the Colorado beetle. Müller's discovery gave medical workers a powerful tool in preventing a host of infectious diseases carried by these insects, including bubonic plague, dengue fever, elephantitis, encephalitis, leishmaniasis, malaria, sleeping sickness, typhus, yaws, and yellow fever. The compound was used widely during World War II (1939–1945) to control insects in both military and

DICHLORODIPHENYLTRICHLOROETHANE



Dichlorodiphenyltrichloroethane (DDT). Black atoms are carbon; green atoms are chlorine; white atoms are hydrogen. Bonds in the benzene rings are represented by the double striped sticks. White sticks show single bonds.

PUBLISHERS RESOURCE GROUP

civilian facilities. It was credited for reducing the spread of what is probably the world's most common infectious disease, malaria. Once common throughout the world, malaria was essentially eliminated in Europe and North America as a result of the use of DDT.

In addition to its public health applications, DDT has long been a popular agricultural pesticide. The compound kills many of the insects and other pests that attack crops and reduce the productivity that can be achieved by farmers.

The fate of DDT took a dramatic turn in the 1960s, however, largely as the result of a single book, *Silent Spring*, written by American biologist and environmentalist Rachel Carson (1907-1964). The book described a situation in which the use of pesticides such as DDT might become so common that the chemical begins to kill off other organisms, such as birds, as well as insects. Carson's book came at a time when scientific evidence of the environmental harm done by pes-

ticides was just becoming available. The work was soon adopted by the growing environmental movement as a well-documented warning of what might happen if the uncontrolled use of pesticides did not end. People began to worry about the ill effects of DDT and other pesticides not only on the environment, but also on human health.

In 1972, the U.S. Environmental Protection Agency announced that DDT could no longer be used in the United States because of the risk it posed for human health and the environment. A number of other developed nations soon followed the EPA's action. In 2001, a conference sponsored by the United Nations, called the Stockholm Convention on Persistent Organic Pollutants, adopted a statement calling for the eventual elimination of a dozen pesticides, known as "the dirty dozen," including DDT. Because of its effectiveness against malaria and other infectious diseases, DDT was given a waiver for use by public health officials in nations where the compound had not been banned.

DDT is also known by the following names: 1,1,1-trichloro-2,2-bis-(p-chlorophenyl)ethane; 1,1'-(2,2,2-trichloroethylidene)-bis (4-chlorobenzene); 1,1,1-trichloro-2,2-di(4-chlorophenyl) ethane; 1,1,1-trichloro-2,2-di(*p*-chlorophenyl) ethane; 1,1-bis(4-chlorophenyl)-2,2,2-trichlorethane; 1,1-bis(p-chlorophenyl)-2,2,2-trichlorethane; 2,2,2-trichloro-1,1-bis (4-chlorophenyl) ethane; and 4,4'-dichlorodiphenyltrichloroethane.

HOW IT IS MADE

DDT is prepared commercially in the reaction between chloral (CCl_3CHO) and chlorobenzene ($\text{C}_6\text{H}_5\text{Cl}$): $\text{CCl}_3\text{CHO} + 2\text{C}_6\text{H}_5\text{Cl} \rightarrow (\text{ClC}_6\text{H}_4)_2\text{CHCCl}_3$.

COMMON USES AND POTENTIAL HAZARDS

DDT is no longer used in the United States and other nations where it has been banned. It is still used in a number of developing nations for the control of infectious diseases spread by insects that can be killed by DDT. Malaria is the most important of these diseases.

The health hazards posed by DDT are still the subject of intense debate among agricultural scientists, public health experts, and other authorities. Abundant evidence exists to show that DDT can have a variety of harmful effects on

Interesting Facts

- Peak use of DDT in the United States occurred in 1958 when about 35 million kilograms (80 million pounds) of the compound were sprayed on U.S. farmlands.
- Extensive data exists to support the advantages of using DDT against malaria. For example, there were more than 2.8 million cases of malaria in Sir Lanka in 1948, before the introduction of DDT. Six years later, after the compound had been put into use in the country, the number of malaria cases dropped to 17. Five years later, after the use of DDT had been discontinued, the number of malaria cases rose to 2.5 million.
- Countries that have banned the use of DDT include Australia (1967), Sweden (1970), Cuba (1970), Germany (1974), Poland (1976), the United Kingdom (1984), Chile (1985), South Korea (1986), Switzerland (1986), and Canada (1989).
- DDT is still manufactured legally in only three countries: China, India, and Indonesia. It may also be produced, although illegally, in Mexico.
- DDT often occurs in combination with two other related compounds, dichlorodiphenylchloroethane (DDD) and dichlorodiphenylchloroethylene (DDE).

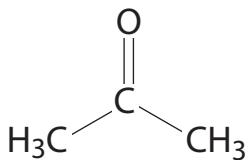
animals, ranging from marine algae to bald eagles. The compound may also have serious health effects on humans, including mild to serious damage to the liver, kidneys, and central nervous system. DDT has also been classified as a probable carcinogen by the U.S. Environmental Protection Agency although no definite evidence for that classification is yet available.

On the other hand, critics have pointed out that little scientific evidence exists for the belief that DDT is a serious health threat to humans. In one experiment, humans took capsules containing small amounts of DDT every day for 18 months and experienced no measurable health effects as a result. Some observers argue that the benefits gained from

using DDT as an agricultural and public health pesticide—benefits such as the elimination of diseases like malaria—far exceed the risks posed by the chemical to human health and the environment.

FOR FURTHER INFORMATION

- Carson, Rachel. *Silent Spring*. Boston: Houghton Mifflin, 1962.
- “DDT.” National Safety Council.
<http://www.nsc.org/library/chemical/ddt.htm> (accessed on December 29, 2005).
- “DDT (dichlorodiphenyltrichloroethane).” Extoxnet.
<http://ace.orst.edu/cgi bin/mfs/o1/pips/ddt.htm> (accessed on December 29, 2005).
- Dunlap, Thomas. *DDT: Scientists, Citizens, and Public Policy*. Princeton, NJ: Princeton University Press, 1983.
- Edwards, J. Gordon, and Steven Milloy. “100 Things You Should Know about DDT.” JunkScience.com.
<http://www.junkscience.com/ddtfaq.htm> (accessed on December 29, 2005).
- Karaim, Reed. “Not So Fast with the DDT: Rachel Carson’s Warnings Still Apply.” *American Scholar* (June 2005): 53–60.
- McGinn, Anne Platt. “Malaria, Mosquitoes, and DDT: The Toxic War against a Global Disease.” *World Watch* (May 1, 2002): 10–17.
- “ToxFAQs™ for DDT, DDE, and DDD.” Agency for Toxic Substances and Disease Registry.
<http://www.atsdr.cdc.gov/tfacts35.html> (accessed on December 29, 2005).



OTHER NAMES:
2-propanone; acetone

FORMULA:
 CH_3COCH_3

ELEMENTS:
Carbon, hydrogen,
oxygen

COMPOUND TYPE:
Ketone (organic)

STATE:
Liquid

MOLECULAR WEIGHT:
58.08 g/mol

MELTING POINT:
−94.7°C (−138°F)

BOILING POINT:
56.05°C (132.9°F)

SOLUBILITY:
Miscible with water,
alcohol, ether,
benzene, and
chloroform

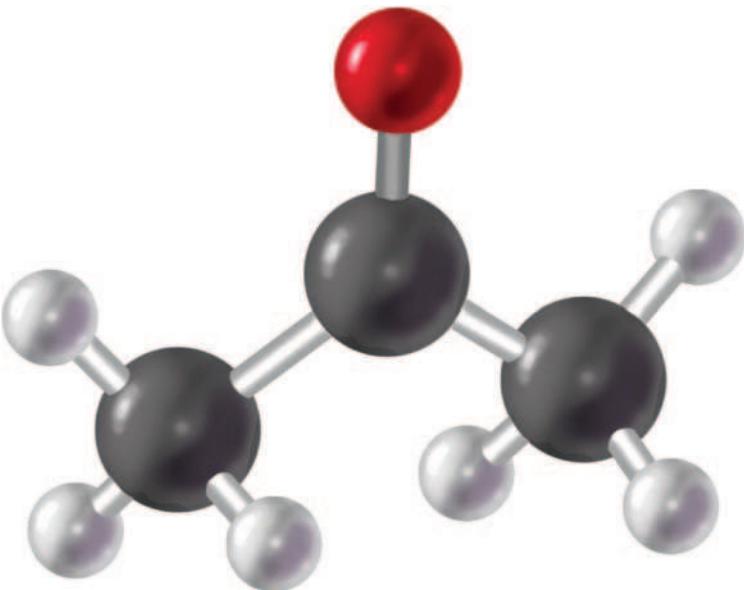
KEY FACTS

OVERVIEW

Dimethyl ketone (DYE-meth-ēl KEY-tone) is a clear, colorless, highly volatile and highly flammable liquid with a characteristic sweet odor and taste. The compound is almost universally known in chemistry laboratories and industrial applications by its common name of acetone.

Acetone was apparently first prepared in 1610 by the French alchemist Jean Béguin (c. 1550–c. 1650). Béguin obtained acetone by heating lead acetate (also known as Saturn's salt) to a high temperature. He obtained a sweet-smelling, very flammable liquid that he named "burning spirit of Saturn." One of the first uses to which the substance was put was as a solvent in the extraction of the active constituents of opium. In 1833, the French chemist Antoine Bussy (1794–1882) gave the compound its modern name of acetone. The correct chemical formula for acetone was determined independently in 1832 by the French chemist Jean Baptiste André Dumas (1800–1884) and the German chemist Justus von Liebig (1803–1873).

Dimethyl ketone. Red atom is oxygen; white atoms are hydrogen; and black atoms are carbon. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP



HOW IT IS MADE

Most of the acetone produced today is made by one of four methods:

- In the Hock process, cumene [$C_6H_5CH(CH_3)_2$] is first oxidized to produce cumene hydroperoxide [$C_6H_5C(CH_3)_2COOH$], which is then reduced to produce acetone and phenol (C_6H_5OH); or
- Isopropyl alcohol (2-propanol; $CH_3CHOHCH_3$) is oxidized over a catalyst to obtain acetone; or
- Butane (C_4H_{10}) is oxidized to obtain acetone; or
- Acetone is obtained as a by-product of the manufacture of glycerol [$C_3H_5(OH)_3$].

COMMON USES AND POTENTIAL HAZARDS

Acetone's primary applications are based on its ability to dissolve such a wide array of organic substances. It is used as a solvent for paints, varnishes, lacquers, inks, glues, rubber cements, fats, oils, waxes, and various types of rubber and plastics. It is perhaps best known to the average person as the primary ingredient in nail polish remover. The largest single use of the compound is as a raw material in the manufacture of other organic chemicals, such as chloroform,

Words to Know

MISCIBLE able to be mixed; especially applies to the mixing of one liquid with another.

VOLATILE A term describing a liquid that can be easily changed to gas.

acetic acid, iodoform, bromoform, isoprene, rayon, and photographic film. It also finds application in storing acetylene gas (because it absorbs up to 24 times its own weight of the gas), to clean and dry chemical equipment and electronic parts, and for the extraction of components of plant and animal tissues.

The primary safety concern about acetone is its extreme flammability. Workers who handle the compound must use great care to prevent its coming into contact or even being in the vicinity of open flames. Under the proper conditions, acetone is also explosive. Exposure of the skin, eyes, and respiratory system to acetone may produce mild symptoms, such as dizziness, headaches, and disorientation and irritation of the eyes and skin. Such conditions are rare, however, and no long-term health effects of the compound have as yet been discovered.

FOR FURTHER INFORMATION

“Acetone.” International Labour Organization.

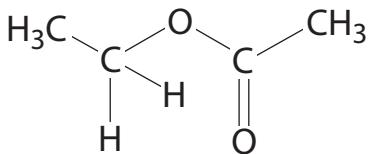
http://www.ilo.org/public/english/protection/safework/cis/products/icsc/dtasht/_icscoo/icscoo87.htm (accessed on October 7, 2005).

“Safety (MSDS) data for propanone.” Physical and Theoretical Chemistry Lab Safety.

<http://ptcl.chem.ox.ac.uk/MSDS/PR/propanone.html> (accessed on October 7, 2005).

“Toxicological Review of Acetone.” Environmental Protection Agency.

http://www.epa.gov/iris/toxreviews/0128_tr.pdf (accessed on October 7, 2005).

**OTHER NAMES:**

Acetic acid; ethyl ester; acetic ether

FORMULA:

CH₃COOC₂H₅

ELEMENTS:

Carbon, hydrogen, oxygen

COMPOUND TYPE:

Ester (organic)

STATE:

Liquid

MOLECULAR WEIGHT:

88.11 g/mol

MELTING POINT:

-83.8°C (-118°F)

BOILING POINT:

77.11°C (170.8°F)

SOLUBILITY:

Slightly soluble in water; miscible with alcohol, ether, and benzene

K E Y F A C T S**OVERVIEW**

Ethyl acetate (ETH-uhl ASS-uh-tate) is a clear, colorless, volatile, flammable liquid with a pleasant fruity odor. Its appealing odor and fruity taste (in dilute solutions) explains one of its primary uses: as an additive in foods and drugs to improve their flavor.

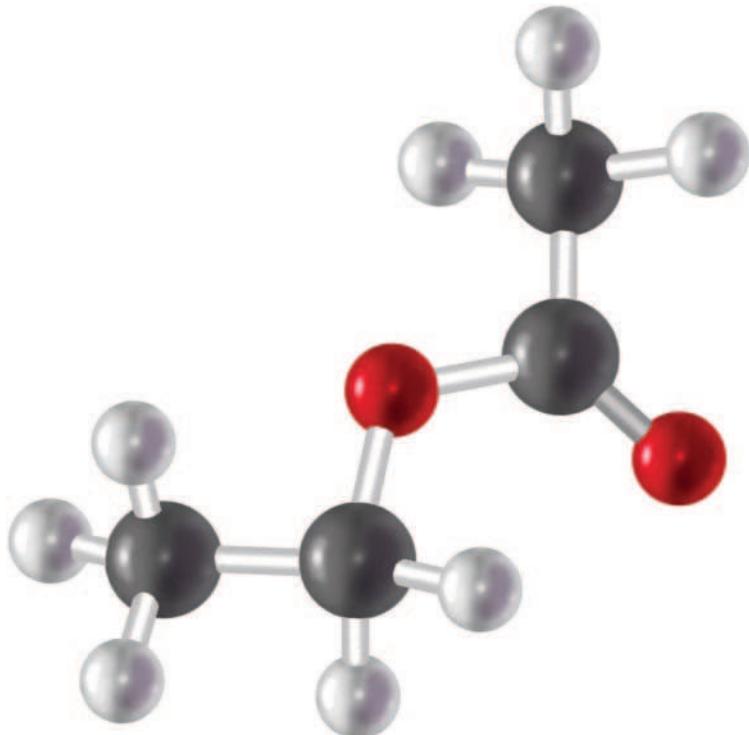
HOW IT IS MADE

Ethyl acetate occurs naturally in fruits, where it is responsible for the pleasant odor and taste of the fruit. It is also found in yeasts and sugar cane. The compound is made synthetically by reacting acetic acid (CH₃COOH) with ethanol (ethyl alcohol; CH₃CH₂OH) in the presence of a sulfuric acid catalyst.

COMMON USES AND POTENTIAL HAZARDS

Ethyl acetate's primary use is as a solvent in a variety of commercial and industrial applications. About 65 percent

Ethyl acetate. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP



of all ethyl acetate produced is used as a solvent in lacquers, varnishes, shellacs, inks, coatings for airplanes, and other types of coatings. Another 12 percent of the ethyl acetate produced is used as a solvent for various types of plastics. The remaining amount of ethyl acetate is used in

Interesting Facts

- Ethyl acetate is a byproduct of the fermentation of grapes in the wine-making process. The compound is also produced as wine ages in barrels and bottles. A moderate amount of ethyl

acetate contributes to the pleasant fruity taste and odor of some wines. But in excess, it can give wine a slightly “off” taste and odor.

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

ESTER An organic compound formed in the reaction between an organic acid and an alcohol.

MISCELLY able to be mixed; especially applies to the mixing of one liquid with another.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

the preparation of other organic compounds; as a flavor-enhancing additive for foods and pharmaceuticals; in the manufacture of smokeless powder, leather products, and photographic films and plates; and for the cleaning of textiles.

The primary safety concern with regard to ethyl acetate is its flammability. Both the liquid and its vapors ignite at temperatures above 200°C (392°F). Exposure to ethyl acetate vapors can also irritate the eyes, nose, and respiratory system. There is no evidence that normal amounts of ethyl acetate are carcinogenic or toxic to humans.

FOR FURTHER INFORMATION

“Ethyl Acetate.” J. T. Baker.

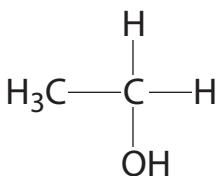
<http://www.jtbaker.com/msds/englishhtml/e2850.htm> (accessed on October 7, 2005).

“Ethyl Acetate.” National Institute of Standards and Technology.
<http://webbook.nist.gov/cgi/cbook.cgi?Name=ethyl+acetate&Units=SI> (accessed on October 7, 2005).

“Ethyl Acetate.” Spectrum Laboratories.

<http://www.speclab.com/compound/c141786.htm> (accessed on October 7, 2005).

See Also Acetic Acid; Ethyl Alcohol



OTHER NAMES:
Ethanol; grain alcohol; alcohol; ethyl hydrate
FORMULA:
$\text{CH}_3\text{CH}_2\text{OH}$
ELEMENTS:
Carbon, hydrogen, oxygen
COMPOUND TYPE:
Alcohol (organic)
STATE:
Liquid
MOLECULAR WEIGHT:
46.07 g/mol
MELTING POINT:
-114.14°C (-173.45°F)
BOILING POINT:
78.29°C (172.9°F)
SOLUBILITY:
Miscible with water, ether, acetone, and most common organic solvents

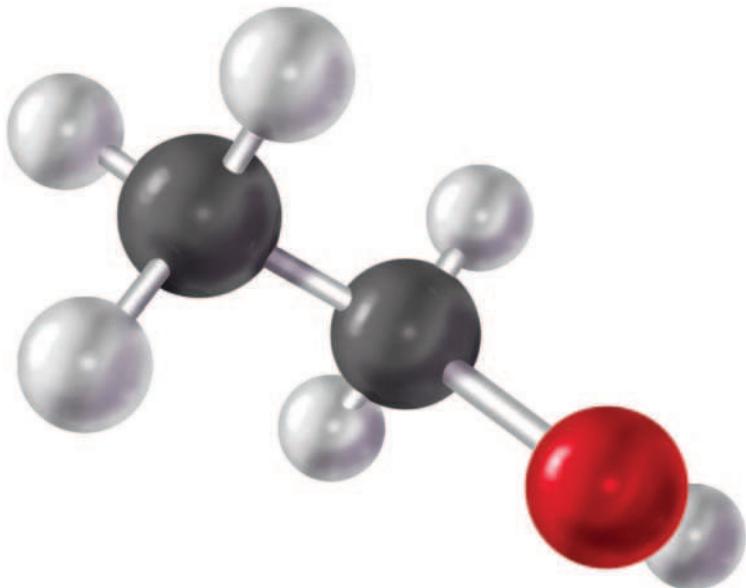
KEY FACTS

Ethyl Alcohol

OVERVIEW

Ethyl alcohol (ETH-uhl AL-ko-hol) is a clear, colorless, flammable liquid with a sharp, burning taste and a pleasant, wine-like odor. It is one of the first chemical substances discovered and used by humans. Ceramic jugs apparently designed to hold beer have been dated to the Neolithic Period, about 10,000 BCE. Some scholars suggest that humans may have learned how to make beer and incorporated it into their daily diets even before they made and used bread. The making and use of wine is a clear theme in Egyptian pictographs dating to the fourth millennium BCE. There probably does not exist a human culture today in which alcohol consumption does not occur. Today, beverages with alcohol content ranging as low as two to five percent ("near beer" and beer) to as high as 50 percent (some forms of vodka) are known and consumed by humans. In spite of its widespread use as a beverage, ethyl alcohol has a number of commercial and industrial uses that account for more than 90 percent of all the compound produced in the United States.

Ethyl alcohol. Red atom is oxygen; white atoms are hydrogen; and black atoms are carbon. PUBLISHERS RESOURCE GROUP



HOW IT IS MADE

Ethyl alcohol is made in one of two ways: naturally, through the process of fermentation, or synthetically, beginning with compounds found in petroleum. Until the beginning of World War II, more than 90 percent of all ethyl alcohol produced in the United States and other developed nations was made by fermentation. Waste syrup left over from the production of sugar from sugar cane was treated with enzymes at temperature of 20°C to 38°C (68°F to 100°F) for 28 to 72 hours. Under these conditions, about 90 percent of the syrup is converted to ethyl alcohol.

Over time, synthetic methods for the production of ethyl alcohol were developed. In one such method, ethylene (ethene; $\text{CH}_2=\text{CH}_2$) is treated with sulfuric acid and water to obtain ethyl alcohol. That method was popular during the 1950s and 1960s. Then, a new method for making the compound was invented. In that process, ethylene and water are heated together at high temperatures [300°C to 400°C (570°F to 750°F)] and high pressures [1,000 pounds per square inch (6.9 megaPascals)] over a catalyst of phosphoric acid (H_3PO_4). The efficiency of this method is greater than

Interesting Facts

- All members of the alcohol family of organic compounds (such as methyl alcohol, ethyl alcohol, and isopropyl alcohol) are toxic to some extent. Only ethyl alcohol is safe to drink in relatively small quantities. Even then, a blood alcohol concentration of less than 5 percent can result in death.
- The concentration of alcohol in a beverage is often expressed as a “proof” number. The proof number of a beverage is twice that of its alcohol concentration. Thus, a beverage that is 80 proof has an alcohol concentration of 40 percent.

the older method, and there are fewer environmental consequences from making ethyl alcohol by this process.

As of 2003, about 94 percent of all ethyl alcohol was produced by fermentation. The remainder was produced by the phosphoric acid method.

COMMON USES AND POTENTIAL HAZARDS

In 2005, 10,500 million liters (2,790 million gallons) of ethyl alcohol were produced by fermentation methods. Of that amount, 92 percent was used as a fuel or an additive in fuels. Many experts suggest that consumers use a mixture of gasoline (90 percent) and ethyl alcohol (10 percent) called gasohol as a vehicle fuel because it burns more completely and releases fewer harmful byproducts to the environment. Although gasohol has not yet become very popular in the United States, it is widely used in some other parts of the world, most notably, in Brazil.

Of the remaining 8 percent of ethyl alcohol produced by fermentation, half was used in industrial operations, as a solvent or intermediary in the preparation of other chemical

compounds; and half was used in the production of alcoholic beverages.

In 2005, about 650 million liters (170 million gallons) of ethyl alcohol were produced by the phosphoric acid method. Of that amount, 60 percent was used for industrial solvents in the manufacture of toiletries and cosmetics, coatings and inks, detergents and household cleaners, pharmaceuticals, and other products. The remaining 40 percent was used in the preparation of other chemical compounds, including ethyl acrylate, vinegar, ethylamines, ethyl acetate, glycol ethers, and miscellaneous materials.

Ethyl alcohol commonly occurs in one of three general forms. Absolute alcohol is ethyl alcohol that contains less than 1 percent impurities, such as water. Absolute alcohol is very difficult to make because ethyl alcohol will absorb water from the atmosphere or any other source that is available. The ethyl alcohol used in fuels and almost all industrial operations is a mixture of 95 percent ethyl alcohol and 5 percent water. Both absolute and 95 percent ethyl alcohol are extremely toxic. Ingestion of even very small amounts of either liquid has serious health effects that may include death.

The alcohol with which most people commonly come into contact is ethyl alcohol mixed with water in alcoholic beverages, such as beer, wine, gin, vodka, rum, or bourbon. In such beverages, the concentration of ethyl alcohol ranges from a few percent to 50 percent.

The effects produced by ethyl alcohol on the human body depend on the type of beverage consumed and the time taken for consumption. Drinking a 5-percent beer over an hour has a very different effect on the body than drinking a 50-percent vodka in five minutes.

Ethyl alcohol is a central nervous system depressant. After ingestion, it passes through a person's stomach and the small intestine, where it is absorbed rapidly into the bloodstream. It then travels throughout the body, interfering with the normal functioning of the nervous system and producing symptoms such as drowsiness, slurred speech, blurred vision, unsteady gait, impaired judgment, and reduced reaction time. With greater concentrations of alcohol in the blood, these symptoms may become more severe, resulting in coma and death.

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

MISCIBLE able to be mixed; especially applies to the mixing of one liquid with another.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

FOR FURTHER INFORMATION

“Alcohol.” Erowid.

<http://www.erowid.org/chemicals/alcohol/alcohol.shtml> (accessed on October 7, 2005).

“Alcohol: What You Don’t Know Can Harm You.” National Institute on Alcohol Abuse and Alcoholism.

http://pubs.niaaa.nih.gov/publications/WhatUDontKnow_HTML/dontknow.htm (accessed on October 7, 2005).

Boggan, William. “Alcohol and You.”

<http://chemcases.com/alcohol/> (accessed on October 7, 2005).

“Chemical of the Week: Ethanol.”

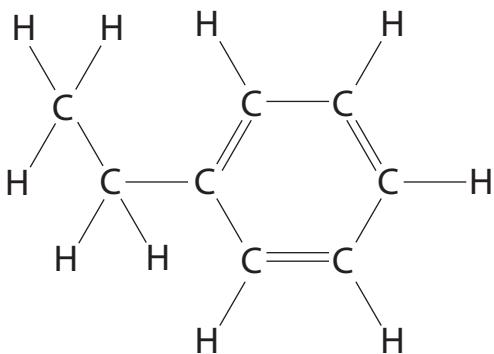
<http://scifun.chem.wisc.edu/chemweek/ethanol/ethanol.html> (accessed on October 7, 2005).

“Ethanol.”

<http://www.ucc.ie/ucc/depts/chem/dolchem/html/comp/ethanol.html> (accessed on October 7, 2005).

“How Alcohol Works.” How Stuff Works.

<http://science.howstuffworks.com/alcohol.htm> (accessed on October 7, 2005).



OTHER NAMES:	Phenylethane; ethylbenzol
FORMULA:	$C_6H_5C_2H_5$
ELEMENTS:	Carbon, hydrogen
COMPOUND TYPE:	Aromatic hydrocarbon (organic)
STATE:	Liquid
MOLECULAR WEIGHT:	106.16 g/mol
MELTING POINT:	-94.96°C (-138.9°F)
BOILING POINT:	136.19°C (277.14°F)
SOLUBILITY:	Immiscible with water; miscible with ethyl alcohol and methyl alcohol

KEY FACTS

Ethylbenzene

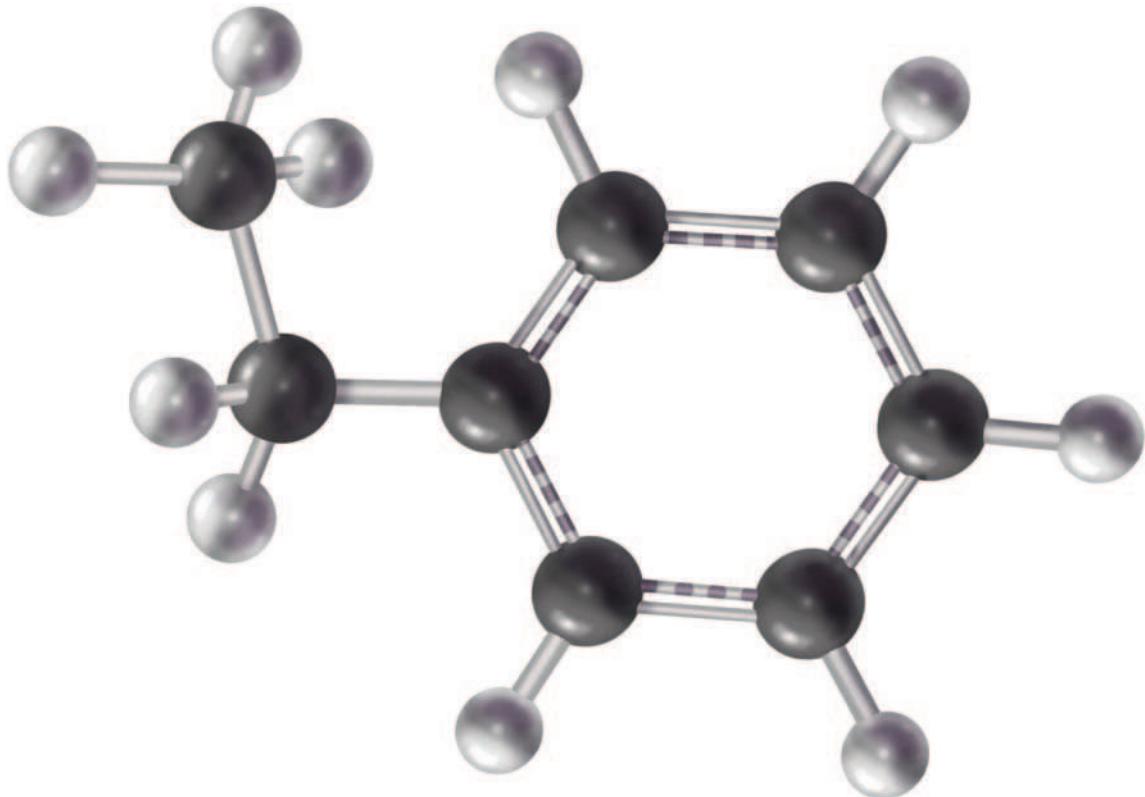
OVERVIEW

Ethylbenzene (eth-il-BEN-zeen) is a colorless flammable liquid with a pleasant aromatic odor. It is an aromatic hydrocarbon, that is, a compound consisting of carbon and hydrogen only with a molecular structure similar to that of benzene (C_6H_6). In 2004 it ranked fifteenth among chemicals produced in the United States. Its primary use is in the manufacture of another aromatic hydrocarbon, styrene ($C_6H_5CH=CH_2$), widely used to make a number of polymers, such as polystyrene, styrene-butadiene latex, SBR rubber, and ABS rubber.

HOW IT IS MADE

Ethylbenzene occurs to some extent as a component of petroleum. It can be extracted from petroleum by fractional distillation, the process by which individual components of petroleum are separated from each other by heating in a distilling tower. Ethylbenzene can also be made synthetically

ETHYLBENZENE



Ethylbenzene. Black atoms are carbon; white atoms are hydrogen. Bonds in the benzene ring are represented by the double striped sticks. White sticks show single bonds.

PUBLISHERS RESOURCE GROUP

by reacting benzene with ethene (ethylene; $\text{CH}_2=\text{CH}_2$) over a catalyst of aluminum chloride (AlCl_3): $\text{C}_6\text{H}_6 + \text{CH}_2=\text{CH}_2 \rightarrow \text{C}_6\text{H}_5\text{C}_2\text{H}_5$.

COMMON USES AND POTENTIAL HAZARDS

More than 99 percent of the ethylbenzene made is used for a single purpose—the production of styrene. Styrene is a very important industrial chemical, ranking seventeenth among all chemicals produced in the United States in 2004. It is used to make a number of important and popular polymers, the best known of which may be polystyrene. Much smaller amounts of ethylbenzene are used in solvents or as additives to a variety of products. Some products that contain ethylbenzene include synthetic rubber, gasoline and other fuels, paints and varnishes, inks, carpet glues, tobacco products, and insecticides.

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

IMMISCIBLE Does not mix with another liquid.

MISCELLY Able to be mixed; especially applies to the mixing of one liquid with another.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

The concentration of ethylbenzene in consumer products is so low that it probably poses no threat to human health or to the environment. It may be a problem, however, when it leaks from industrial or chemical plants into the soil and becomes part of the groundwater. In such cases, it may be consumed by humans and other animals, or it may evaporate into the air, where it may be breathed by humans and other animals. Ethylbenzene has been found in about half (731 of 1467) of all the sites surveyed for pollutants by the U.S. Environmental Protection Agency, although concentrations are so low in most cases as to pose no threat to human health or the environment.

Ethylbenzene is an irritant to the skin, eyes, and respiratory system. In small quantities, it may cause dizziness, tightness in the chest, and burning of the eyes. In larger doses, it may cause narcotic effects, producing drowsiness and disorientation. The greatest safety concerns about ethylbenzene relate to its combustibility. Since it is more dense than air, it tends to settle to the ground and travel to any source of fire that may be near by.

FOR FURTHER INFORMATION

"Consumer Factsheet on: Ethylbenzene." U.S. Environmental Protection Agency.

http://www.epa.gov/safewater/contaminants/dw_contamfs/ethylben.html (accessed on December 22, 2005).

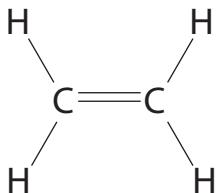
"Ethylbenzene." Australian Government. Department of the Environment and Heritage.

http://www.npi.gov.au/database/substance_info/profiles/40.html (accessed on December 22, 2005).

ETHYLBENZENE

“ToxFAQs™ for Ethylbenzene.” Agency for Toxic Substances and Disease Registry.
<http://www.atsdr.cdc.gov/tfacts110.html> (accessed on December 22, 2005).

See Also Polystyrene; Poly(Styrene-Butadiene-Styrene); Styrene



OTHER NAMES:

Ethene; bicarburetted hydrogen; olefiant gas

FORMULA:

$\text{CH}_2=\text{CH}_2$

ELEMENTS:

Carbon, hydrogen

COMPOUND TYPE:

Alkene; unsaturated hydrocarbon (organic)

STATE:

Gas

MOLECULAR WEIGHT:

28.05 g/mol

MELTING POINT:

-169.15°C (-272.47°F)

BOILING POINT:

-103.77°C
(-154.79°F)

SOLUBILITY:

Insoluble in water; slightly soluble in ethyl alcohol, benzene, and acetone; soluble in ether

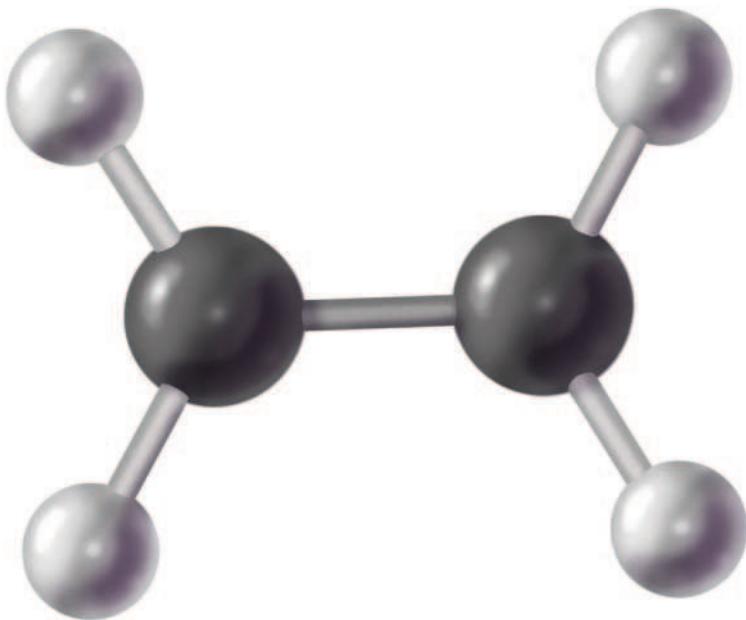
**K
E
Y
F
A
C
T
S**

Ethylene

OVERVIEW

Ethylene (ETH-ih-leen) is a colorless, flammable gas with a sweet odor and taste. It is the simplest alkene. Alkenes are hydrocarbons that contain one or more double bonds. Ethylene was first prepared in 1794 by a group of Dutch chemists including J. R. Deiman, A. Paets van Troostwyk, N. Bondt, and A. Lauwerenburgh. They treated ethanol (ethyl alcohol; $\text{C}_2\text{H}_5\text{OH}$) with concentrated sulfuric acid (H_2SO_4) and obtained ethylene, although they were incorrect in believing that the compound also contained oxygen.

Ethylene occurs naturally in petroleum and natural gas, but only to a very small percentage. It also occurs naturally in plants where it functions as a hormone and has a number of important effects on the growth and development of plants. These effects have been used for thousands of years, although the chemical mechanism involved was not understood. For example, the ancient Chinese are said to have burned incense in closed containers in order to facilitate the ripening of pears. Although they were certainly not



Ethylene. Black atoms are carbon; white atoms are hydrogen. The gray stick between the two carbon atoms is a double bond. White sticks show single bonds. PUBLISHERS RESOURCE GROUP

aware of the fact, the ripening effect was probably a result of ethylene gas released during combustion of the incense.

The hormonal effects of ethylene were first identified in 1901 by Russian chemist Dmitri N. Neljubow. For some time, scientists had observed that leaks from the gas lamps used to light city streets caused plants to grow old more rapidly than normal, often with strange changes in the structure of their leaves and stems. Neljubow was able to show that this effect was caused by ethylene in the lamp gas. Scientists have since learned a great deal about the production of ethylene in plants and the effects produced by the gas on the growth and development of those plants. For example, scientists discovered that ethylene is synthesized in germinating seeds, nodes of stems, and the tissue of ripening fruits. Its production is increased by flooding of the plant's roots; by drying of the soil; in response to environmental stress, such as attack by pests; by aging of the plant; and by inadequate amounts of minerals in the soil. They have also learned that ethylene increases the rate at which leaves and flowers age; promotes germination of seeds and the growth of root hairs; stimulates the ripening of fruit and flowers; and increases a plant's resistance to disease and physical damage.

HOW IT IS MADE

One reference on hydrocarbons lists more than 500 methods for making ethylene. From a practical standpoint, only a handful of those methods are important. The most common method of preparation involves the thermal or catalytic cracking of hydrocarbons. The term *cracking* refers to the process by which hydrocarbons from petroleum are broken down into simpler molecules. That process is usually accomplished by heating the petroleum for short periods of time at high temperatures (*thermal cracking*) or over a catalyst (*catalytic cracking*). In the cracking process, hydrocarbons with 10, 15, 20, or more carbon atoms are broken down to produce hydrocarbons with two, three, four, or some other small number of carbons. Ethylene is one of the usual products of cracking. It can be separated from the other products of cracking because it escapes from the reaction mixture as a gas.

Ethylene can also be produced from synthesis gas. Synthesis gas is the term used for various mixtures of gases produced when steam (with or without additional oxygen) is passed over hot coal. The steam and coal react to produce a rich mixture of hydrocarbons, a mixture that usually includes ethylene. Finally, ethylene can be produced in small quantities in the laboratory by the method first used by the Dutch chemists in 1794, namely by reacting ethanol with concentrated sulfuric acid.

COMMON USES AND POTENTIAL HAZARDS

In 2004, U.S. chemical manufacturers produced about 26.7 million metric tons (29.4 million short tons) of ethylene, making it the third most important chemical produced in the country in terms of volume. Over 90 percent of that ethylene was used for the production of other chemical compounds, the most important of which were polymers and related compounds, such as three major kinds of polyethylene (high density (HDP), low density (LDP), and linear low density polyethylene (LLDP)), polypropylene, ethylene dichloride, ethylene oxide, styrene, ethylbenzene, vinyl acetate, vinyl chloride, ethylene glycol, polystyrene, polyvinyl chloride (PVC), trichloroethylene, styrene-butadiene rubber (SBR), and ethyl alcohol.

Interesting Facts

- The highest rate of ethylene production yet measured in plants is that produced by the fading blossoms of the Vanda orchid, at 3.4 microliters per gram of flower per hour. The Vanda orchid is extensively used in the manufacture of Hawaiian leis.
- Air in rural areas typically contains about 5 parts per billion (ppb) of ethylene, while urban air generally contains about twenty times that amount.
- Ethylene is produced in plants in a complex series of reactions known as the Yang cycle (after its discoverer, S. F. Yang) that starts with the amino acid methionine.
- Production of ethylene in plants is significantly affected by a number of factors, such as temperature and presence of other gases (especially oxygen and carbon dioxide).

Smaller amounts of ethylene were used in a number of other chemical and industrial applications. These included:

- As a spray to accelerate the ripening of fruit;
- As a refrigerant;
- In oxygen-ethylene torches used in welding and metal-cutting operations;
- As a specialized anesthetic.

Ethylene poses a health hazard primarily because it is highly flammable and a serious explosive risk. It also acts as a narcotic at low concentrations, causing nausea, dizziness, headaches, and loss of muscular coordination. At higher concentrations, it acts as an anesthetic, causing loss of consciousness and insensitivity to pain and other stimuli. These effects tend to be of concern primarily to people who work directly with the gas. The amount of ethylene to which most people are exposed in their daily lives tends to be relatively low.

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

CRACKING The process by which larger hydrocarbons from petroleum

are broken down into simpler molecules.

POLYMER : A compound consisting of very large molecules made of one or two small repeated units called monomers.

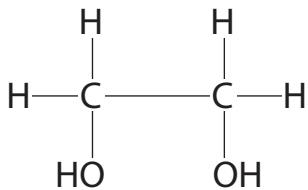
FOR FURTHER INFORMATION

“Ethene (ethylene): Properties, Production & Uses.” Aus e tute.
<http://www.ausetute.com.au/ethene.html> (accessed on December 27, 2005).

“Ethylene.” Laboratory of Postharvest Physiology and Technology, Seoul National University.
<http://plaza.snu.ac.kr/~postharv/p3ethylene.pdf> (accessed on December 27, 2005).

“Ethylene.” National Safety Council.
<http://www.nsc.org/library/chemical/Ethylene.htm> (accessed on December 27, 2005).

See Also Ethyl Alcohol; Ethylene Glycol; Ethylene Oxide; Ethylbenzene; Polyethylene; Polypropylene; Polystyrene; Polyvinyl Chloride



OTHER NAMES:
Ethylene alcohol;
monoethylene glycol

FORMULA:
 $\text{CH}_2\text{OHCH}_2\text{OH}$

ELEMENTS:
Carbon, hydrogen,
oxygen

COMPOUND TYPE:
Dihydric alcohol
(organic)

STATE:
Liquid

MOLECULAR WEIGHT:
62.07 g/mol

MELTING POINT:
−12.69°C (9.16°F)

BOILING POINT:
197.3°C (387.1°F)

SOLUBILITY:
Miscible with
water, alcohol,
acetone, and ether

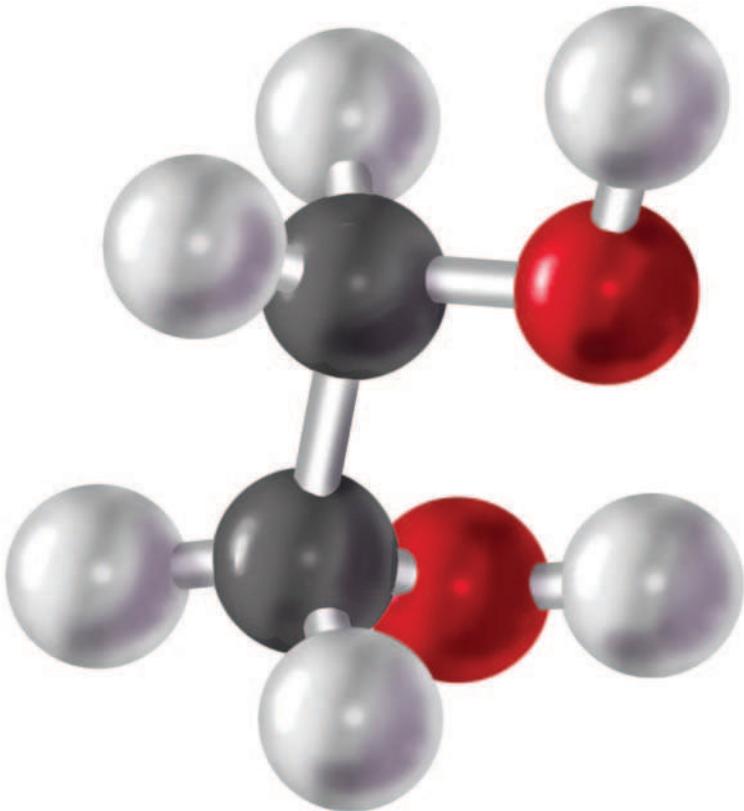
KEY FACTS

OVERVIEW

Ethylene glycol (ETH-uh-leen GLYE-kol) is clear, colorless, syrupy liquid with a sweet taste. One should not attempt to confirm the compound's taste, however, as it is toxic. In recent years, more than 4 billion kilograms (9 billion pounds) of ethylene glycol has been produced in the United States annually. The compound is used primarily as an antifreeze and in the manufacture of a number of important chemical compounds, including polyester fibers, films, bottles, resins, and other materials.

Ethylene glycol was first prepared in 1859 by the French chemist Charles Adolphe Wurtz (1817–1884). Wurtz's discovery did not find an application, however, until the early twentieth century, when the compound was manufactured for use in World War I (1914–1918) in the manufacture of explosives and as a coolant. By the 1930s, a number of uses for the compound had been found, and the chemical industry began producing ethylene glycol in large quantities.

Ethylene glycol. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. PUBLISHERS RESOURCE GROUP



HOW IT IS MADE

The primary method of producing ethylene glycol involves the hydration of ethylene oxide, a ring compound consisting of two methylene (-CH_2) groups and one oxygen atom. Hydration is the process by which water is added to a compound. The hydration of ethylene oxide is conducted at a temperature of about 383°F (195°C) without a catalyst, or at about 50°C to 70°C (122°F to 158°F) with a catalyst, usually a strong acid, either process resulting in a yield of at least 90 percent of ethylene glycol.

Other methods of preparation are also available. For example, the compound can be produced directly from synthesis gas, a mixture of carbon monoxide and hydrogen; or by treating ethylene ($\text{CH}_2=\text{CH}_2$) with oxygen in an acetic acid solution using a catalyst of tellurium oxide or bromide ion.

Interesting Facts

- In 1996, 60 children died in Haiti of ethylene glycol poisoning after drinking cough syrup made in

China that had accidentally been contaminated with the compound.

COMMON USES AND POTENTIAL HAZARDS

One of the first major uses of ethylene glycol was as a radiator coolant in airplanes. The compound actually made possible a change in the design of airplanes. At one time, plain water was used as the coolant in airplane radiators. The faster the airplane flew, the greater the risk that its radiator would boil over. Adding ethylene glycol to the water raised the boiling point of the coolant and allowed airplanes to fly faster with smaller radiators. This change was especially useful in the construction of military airplanes used in combat.

Ethylene glycol is still used extensively as a coolant and antifreeze in cooling systems. It is also used as a deicing fluid for airport runways, cars, and boats. Brake fluids and shock-absorber fluids often contain ethylene glycol as protection against freezing. About 26 percent of all the ethylene glycol made in the United States is used for some kind of cooling or antifreeze application.

The largest single use of ethylene glycol today is in the manufacture of a plastic called polyethylene terephthalate (PET). PET's primary application is in the manufacture of plastic bottles, an application that accounts for about a third of all the ethylene glycol made in the United States. Large amounts of PET are also used in the manufacture of polyester fibers and films. Some additional uses of the compound include:

- As a humectant (a substance that attracts moisture) in keeping some food, tobacco, and industrial products dry;
- As a solvent in some paints and plastics;

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

HUMECTANT A substance that attracts moisture more easily than other substances.

MISCELLY able to be mixed; especially applies to the mixing of one liquid with another.

- In the dyeing of leathers and textiles;
- In the manufacture of printing inks, wood stains, ink for ball-point pens, and adhesives;
- In the production of artificial smoke and fog for theatrical productions;
- As a stabilizer in the soybean-based foam sometimes used to extinguish industrial fires; and
- In the manufacture of specialized types of explosives.

Ethylene glycol poses a number of potential health and safety hazards. It is very flammable and highly toxic. Ingestion of the compound may cause nausea, vomiting, abdominal pain, weakness, convulsions, and cardiac problems. Higher doses can result in severe kidney damage that leads to death.

FOR FURTHER INFORMATION

“Ethylene Glycol.” Environmental Protection Agency.

http://www.epa.gov/ttn/atw/hlthef/ethy_gly.html (accessed on October 7, 2005).

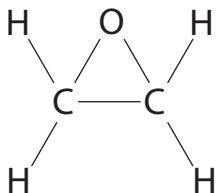
“Ethylene Glycol.” National Safety Council.

<http://www.nsc.org/library/chemical/Ethylen1.htm> (accessed on October 7, 2005).

“Medical Management Guidelines (MMGs) for Ethylene Glycol.”

Agency for Toxic Substances and Disease Registry.

<http://www.atsdr.cdc.gov/MHMI/mmg96.html> (accessed on October 7, 2005).



OTHER NAMES:
Epoxyethane; oxirane

FORMULA:
 $(\text{CH}_2)_2\text{O}$

ELEMENTS:
Carbon, hydrogen,
oxygen

COMPOUND TYPE:
Cyclic ether (organic)

STATE:
Gas

MOLECULAR WEIGHT:
44.05 g/mol

MELTING POINT:
−112.5°C (−170.5°F)

BOILING POINT:
10.6°C (51.1°F)

SOLUBILITY:
Soluble in water,
ethyl alcohol,
acetone, benzene,
and ether

KEY FACTS

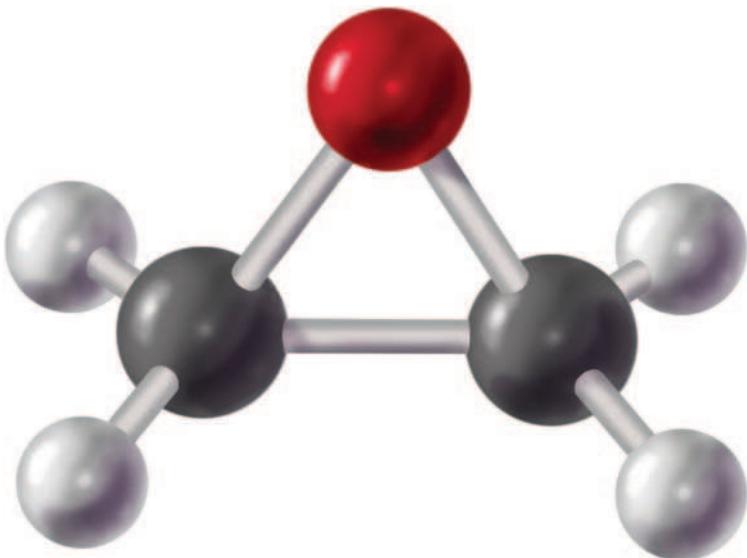
OVERVIEW

Ethylene oxide (ETH-ih-leen OK-side) is a flammable, colorless gas with the odor of ether. The gas is a cyclic compound, consisting of a ring of two carbon atoms and one oxygen atom. Each carbon atom also has two hydrogen atoms attached to it. Ethylene oxide was first prepared in 1859 by French chemist Charles Adolphe Wurtz (1817–1884). Wurtz produced the compound by reacting ethylene chlorhydrin (2-chloroethanol; $\text{ClCH}_2\text{CH}_2\text{OH}$) with an inorganic base (such as sodium hydroxide; NaOH), a process that remained the principle method for preparing the gas for more than a century. After World War II (1939–1945), a method was discovered for the direct oxidation of ethylene gas that is more efficient than the chlorhydrin process.

Ethylene oxide is a very unstable compound that catches fire or explodes readily and must be handled with the greatest care. Nonetheless, it is an important industrial chemical and ranks nineteenth by volume among chemicals produced in the United States. Its primary use is in the manufacture of other organic compounds.

Ethylene Oxide. Black atoms are carbon; white atoms are hydrogen; red atom is oxygen.
White sticks show single bonds.

PUBLISHERS RESOURCE GROUP



HOW IT IS MADE

The chlorhydrin process for making ethylene oxide has been replaced commercially by the direct oxidation of ethylene gas. Oxidation takes place at temperatures of 200°C to 300°C (392°F to 572°F) over a silver catalyst. The formula for this reaction is $2\text{CH}_2=\text{CH}_2 + \text{O}_2 \rightarrow 2\text{CH}_2\text{CH}_2\text{O}$. The yield produced by direct oxidation is slightly less than that produced by the chlohydrin process, but the amount of chlorine wasted by the latter method outweighs the slight difference in efficiency of production.

COMMON USES AND POTENTIAL HAZARDS

The largest single use of ethylene oxide is in the manufacture of ethylene glycol ($\text{CH}_2\text{OHCH}_2\text{OH}$), which itself is used as an antifreeze and raw material for the production of plastics. About 60 percent of all the ethylene oxide produced is used for this purpose. The compound is also used to make higher glycols, such as diethylene and triethylene glycol ($\text{CH}_2\text{OHCH}_2\text{OCH}_2\text{CH}_2\text{OH}$ and $\text{CH}_2\text{OHCH}_2\text{OCH}_2\text{CH}_2\text{OCH}_2\text{CH}_2\text{OH}$). The second most important application of ethylene oxide is in the synthesis of ethoxylates and ethanolamines, substances used in the production of synthetic detergents. These substances act as surfactants in detergents—substances

Interesting Facts

- Ethylene oxide's instability is caused by its unusual three-atom ring structure. A ring with three atoms is a stressed arrangement that breaks apart with even moderate amounts of stress.
- A number of industrial accidents have been caused during the preparation, transportation, storage, and use of ethylene oxide. It is one of the most hazardous of the top twenty-five chemicals produced in the United States.
- In 2004, some 3.77 million metric tons (4.15 million short tons) of ethylene oxide were produced in the United States.

that reduce the surface tension between two materials and improve the lathering ability of a cleaner.

Ethylene oxide also has a number of other important industrial uses, although the quantity used for such purposes is small compared to the uses mentioned above. For example, it is used as a rocket propellant because of its tendency to decompose easily with the release of large amounts of energy. It is also used as a sterilizing medium, particularly for the sterilization of surgical instruments and consumer products, such as spices and cosmetics. The compound is also used as a demulsifier in the petroleum industry. A demulsifier is a material that aids in the separation of the components of complex mixtures, like those handled in the processing of petroleum. Ethylene oxide is also used as a fumigant, which is a gas used to kill insects and other pests.

A number of health hazards are associated with exposure to ethylene oxide. It is a skin, eye, and respiratory system irritant, causing symptoms such as dizziness, nausea, headache, convulsions, blistering of the skin, coughing, tightness of the chest, difficulty in breathing, and blurred vision. Long-term exposure to the gas may produce more serious health consequences, such as damage to the nervous system, muscular weakness, paralysis of the peripheral nerves, impaired thinking, loss of memory, and severe skin irritation. Ethylene

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

ETHER An organic compound in which one oxygen atom is bonded to two carbon atoms: - C - O - C. Ethylene oxide is the simplest cyclic ether.

oxide is believed to be responsible for spontaneous abortions, genetic damage, and the development of some types of cancer, primarily cancer of the blood (leukemia).

FOR FURTHER INFORMATION

Buckles, Carey, et al. *Ethyleneoxide*. 2nd ed. Celanese Ltd., The Dow Chemical Company, Shell Chemical Company, Sunoco, Inc., and Equistar Chemicals, LP. Also available online at <http://www.ethyleneoxide.com/html/introduction.html> (accessed January 3, 2006).

Environment Canada, Health Canada. *Ethylene Oxide*. Ottawa: Environment Canada, 2001.

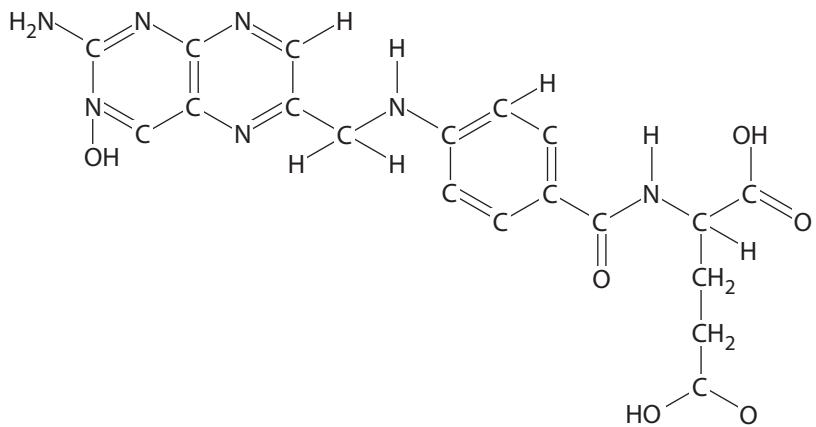
“Ethylene Oxide.” OSHA Fact Sheet.

http://www.osha.gov/OshDoc/data_General_Facts/ethylene_oxide_factsheet.pdf (accessed on December 27, 2005).

“Ethylene Oxide Health and Safety Guide.” IPCS International Programme on Chemical Safety.
<http://www.inchem.org/documents/hsg/hsg/hsgo16.htm>
(accessed on December 27, 2005).

“ToxFAQs™ for Ethylene Oxide.” Agency for Toxic Substances and Disease Registry.
<http://www.atsdr.cdc.gov/tfacts137.html> (accessed on December 27, 2005).

See Also Ethylene; Ethylene Glycol

**OTHER NAMES:**

L-glutamic acid;
vitamin Bc; vitamin
B₉; vitamin M

FORMULA:

C₁₉H₁₉N₇O₆

ELEMENTS:

Carbon, hydrogen,
nitrogen, oxygen

COMPOUND TYPE:

Organic acid

STATE:

Solid

MOLECULAR WEIGHT:

441.40 g/mol

MELTING POINT:

Decomposes at 250°C
(480°F)

BOILING POINT:

Not applicable

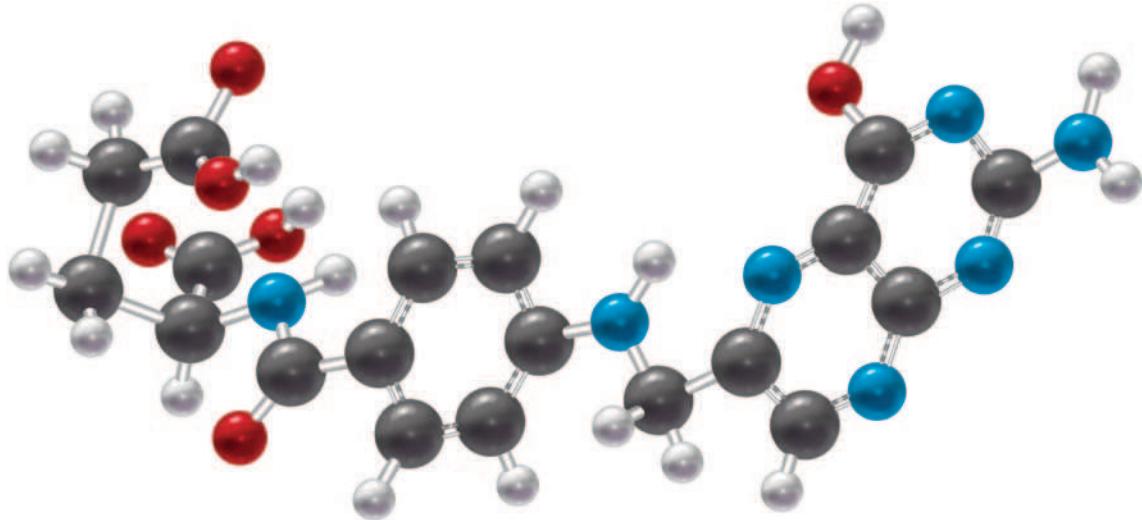
SOLUBILITY:

Very slightly soluble
in water and methyl
alcohol; insoluble in
ethyl alcohol and
acetone

KEY FACTS**OVERVIEW**

Folic acid (FOH-lik AS-id) is a member of the B vitamin group, which is essential for the production of proteins and nucleic acids. In pure form, it is a tasteless, odorless, orange-to-yellow crystalline substance that is destroyed by heat or exposure to light. The compound occurs in three similar forms with comparable biological activity. Only one of its forms, L-pteroylglutamic acid, is made synthetically (in a laboratory). Folic acid is sometimes referred to in its ionic form as folate, which differs from folic acid only in the absence of a single hydrogen atom in its structure. The term folate is also used for a group of compounds structurally similar to folic acid.

Credit for the discovery of the nutritional significance of folic acid is often given to English medical researcher Lucy Wills (1888-1964). In the early 1920s, Wills discovered that anemia in pregnant women could be prevented if they included yeast in their diets. Anemia is a condition in which the blood contains too few red blood cells. Wills located



Folic Acid. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atoms are nitrogen. Striped sticks indicate a benzene ring.

PUBLISHERS RESOURCE GROUP

a specific compound in yeast that produced this effect and called it the “Wills Factor.” At about the same time, other research teams were discovering a similar compound that prevented anemia in monkeys, chicks, and other animals. One team, led by American researcher William J. Darby (1913–2001) called their anti-anemia factor “vitamin M,” (for monkeys) and a third research team discovered a similar factor that prevented anemia in chicks and called it vitamin Bc (for chicks). Folic acid was finally isolated and identified in 1941 by American researcher Henry K. Mitchell (1917–), who suggested the modern name of folic acid for the compound. He chose the name because the compound was abundant in leafy vegetables and the Latin word for leaf is *folium*.

By the 1980s, scientists had produced evidence that the addition of folic acid in the diets of pregnant women can prevent birth defects such as spina bifida, a condition in which a baby’s spinal column fails to close properly while developing inside the mother’s womb. Researchers had learned how to fortify foods with folic acid in the 1970s, but manufacturers did not actually begin to do so until the late 1990s, when the U.S. government began requiring companies to supplement cereals, breads, and other grain-based products with the vitamin.

Interesting Facts

- The recommended daily dose of folic acid for adults is 400 micrograms per day. The best sources of the vitamin are fortified

cereals and grain products, beef liver, black-eyed peas, spinach, avocado (raw), and eggs.

HOW IT IS MADE

The body produces some folic acid and obtains the remainder through food and dietary supplements. In the body, folic acid is produced by bacteria in the large intestine, absorbed in the small intestine, and stored in the liver. The L-pteroylglutamic acid form of folic acid is also produced synthetically.

COMMON USES AND POTENTIAL HAZARDS

The sole use of folic acid is as a nutrient in animal bodies. It is used in the synthesis of methionine, an amino acid used in the formation of proteins and nucleic acids. A deficiency of folic acid can produce various symptoms, including ulcers in the stomach and mouth, slowed growth, and diarrhea. It also results in a medical condition known as megaloblastic anemia, in which a person's body produces red blood cells that are larger than normal.

Adequate amounts of folic acid are especially important in fetal development, during the first eight weeks of life following fertilization. The compound is essential to promote normal development of the fetal nervous system. Folic acid deficiencies in the mother during this period may result in neural tube defects such as spina bifida or anencephaly, a condition in which the fetus' brain and skull fail to develop normally. The U.S. Centers for Disease Control and Prevention recommend that pregnant women take 600 micrograms of folic acid daily to avoid such problems.

Some evidence suggests that folic acid supplements may also reduce the risk of heart disease; strokes; and cervical and colon cancers. Folic acid is generally nontoxic, and side

Words to Know

SYNTHESIS A chemical reaction in which some desired chemical product is made

from simple beginning chemicals, or reactants.

effects associated with its use are very rare. In unusual cases, allergic reactions to the compound have been reported.

FOR FURTHER INFORMATION

“Folate.” PDRhealth.

http://www.pdrhealth.com/drug_info/nmdrugprofiles/nutsupdrugs/fol_0110.shtml (accessed on October 10, 2005).

Folate (Folacin, Folic Acid). Ohio State University Extension Fact Sheet.

<http://ohioline.osu.edu/hygfact/5000/5553.html> (accessed on October 10, 2005).

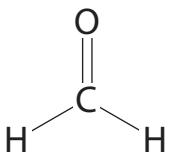
“Folic Acid.” International Programme on Chemical Safety.

<http://www.inchem.org/documents/pims/pharm/folicaci.htm> (accessed on October 10, 2005).

“Folic Acid Fortification.” U.S. Food and Drug Administration.

http://vm.cfsan.fda.gov/~dms/wh_folic.html (accessed on October 10, 2005).

“Vitamins: The Quest for Just the Right Amount.” *Harvard Health Letter*. (June 2004): 1.

**OTHER NAMES:**

Methanal;
oxomethylene;
oxomethane;
methylene oxide;
formic aldehyde

FORMULA:

HCHO

ELEMENTS:

Carbon, hydrogen,
oxygen

COMPOUND TYPE:
Aldehyde (organic)

STATE:
Gas

MOLECULAR WEIGHT:
30.03 g/mol

MELTING POINT:
−92°C (−130°F)

BOILING POINT:
−19.1°C (−2.38°F)

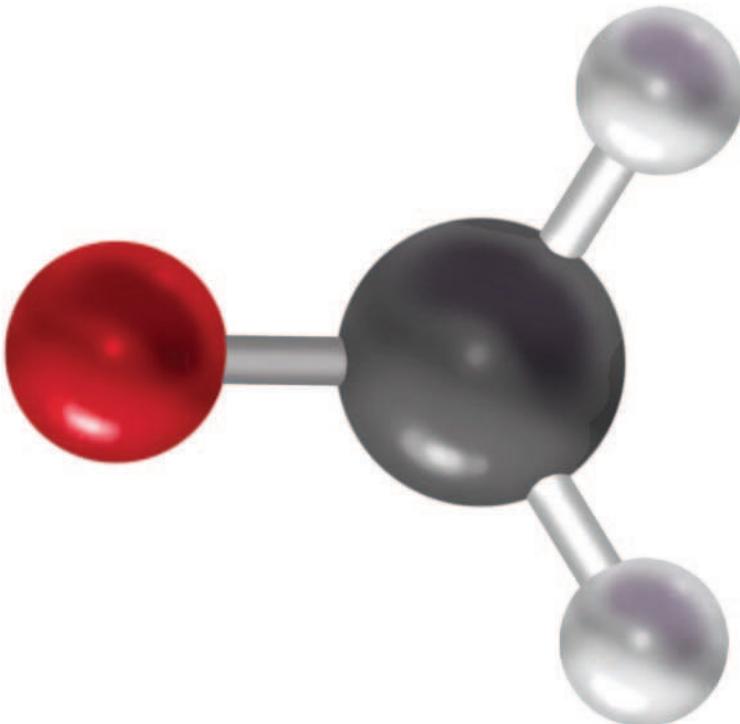
SOLUBILITY:
Very soluble in water,
alcohol, ether, and
benzene

K E Y F A C T S**OVERVIEW**

Formaldehyde (for-MAL-duh-hide) is a colorless, flammable gas with a strong, pungent odor that tends to polymerize readily. Polymerization occurs when individual molecules of formaldehyde combine with each other to make very large molecules called polymers. Over 4 billion kilograms (10 billion pounds) of formaldehyde were produced in the United States in 2004, the vast majority of which was used in the production of plastics and other polymers. To make handling and shipping easier and safer, the compound is usually provided as a 37 percent solution of formaldehyde in water to which has been added an additional 15 percent of methanol (methyl alcohol) to prevent polymerization.

Formaldehyde was first produced accidentally in 1859 by the Russian-French chemist Alexander Mikhailovich Butlerov (1828–1886). It was first synthesized in 1867 by the German chemist August Wilhelm von Hofmann (1818–1892) who was not, however, able to collect the compound in

Formaldehyde. Red atom is oxygen; white atoms are hydrogen; and black atom is carbon. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP



pure form. That step was accomplished by German chemist Friedrich August Kekulé (1829–1896) in 1892.

HOW IT IS MADE

Formaldehyde occurs naturally in the atmosphere at a concentration of about 10 parts per billion (0.000 001%) partly as a by-product of plant and animal metabolism, and partly as a product of the reaction of sunlight with methane (CH_4), a much more abundant component of the air. At such low concentrations, it is not a natural source of the compound for commercial or industrial uses and is produced instead by the oxidation of methanol (methyl alcohol; CH_3OH) or gases extracted from petroleum (such as methane) over a catalyst of silver, copper, or iron with molybdenum oxide.

COMMON USES AND POTENTIAL HAZARDS

By far the most important application of formaldehyde is in the production of polymers and other organic chemicals.

Interesting Facts

- Formaldehyde was one of the first organic compounds to have been discovered in outer space.
- When some vegetables, such as cabbage and brussel sprouts are cooked, they emit small amounts of formaldehyde.

About one-quarter of commercial-use formaldehyde is used each year to make a family of polymers known as urea-formaldehyde resins, which are used to make dinnerware, particle board, fiber board, plywood, flexible foams, and insulation. Another 16 percent goes to the production of phenol-formaldehyde resins, with applications in molded and cast plastics, adhesives and bonding materials, laminating materials, brake linings, chemical equipment, machine housing, and a host of other applications. Smaller amounts of formaldehyde are used to make a variety of important chemicals including 1,4-butanediol, methylene diisocyanate, pentaerythritol, and hexamethylenetetramine. Other applications include use in controlled release fertilizers, in the production of nitroparaffin derivatives, in the treatment of textiles, and in the preservation of biological specimens. The last of these uses is probably well known to biology students; its use depends on the fact that formaldehyde kills most types of bacteria and can be used, therefore, to keep biological materials from decaying.

Formaldehyde poses a number of health hazards to humans and other animals. It may cause difficulty in breathing, headaches, fatigue, and lowered body temperature. At high levels of concentration or over long periods of exposure, formaldehyde can induce coma and death. Chronic exposure to formaldehyde is thought to be carcinogenic, producing tumors in the nose, throat, and respiratory system. People who work in factories where formaldehyde is used are at greatest risk for formaldehyde poisoning.

Formaldehyde is now known to be a potentially serious indoor air pollutant. So many products in a home contain

Words to Know

CARCINOGEN A chemical that causes cancer in humans or other animals.

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

METABOLISM A process that includes all of the chemical reactions that occur in cells

by which fats, carbohydrates, and other compounds are broken down to produce energy and the compounds needed to build new cells and tissues.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

formaldehyde that significant levels of the compound may accumulate in a house. The primary sources of the formaldehyde are pressed wood products such as plywood and particleboard; furnishings; wallpaper; and durable press fabrics.

FOR FURTHER INFORMATION

"About Formaldehyde." Formaldehyde Council.

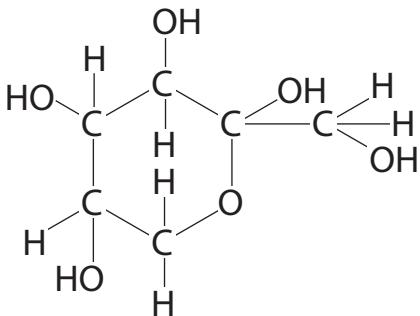
http://www.formaldehyde.org/about_what.html (accessed on October 10, 2005).

"Formaldehyde (Methyl Aldehyde) Fact Sheet." Australian Government; Department of the Environment and Heritage.
http://www.npi.gov.au/database/substance_info/profiles/45.html (accessed on October 10, 2005).

Gullickson, Richard. "Reference Data Sheet on Formaldehyde." Meridian Engineering & Technology.
<http://www.meridianeng.com/formalde.html> (accessed on October 10, 2005).

"An Update on Formaldehyde 1997 Revision." U.S. Consumer Product Safety Commission.
<http://www.epa.gov/iaq/pubs/formald2.html> (accessed on October 10, 2005).

See Also Urea



OTHER NAMES:	D-Fructose; fruit sugar
FORMULA:	$C_6H_{12}O_6$
ELEMENTS:	Carbon, hydrogen, oxygen
COMPOUND TYPE:	Carbohydrate (organic)
STATE:	Solid
MOLECULAR WEIGHT:	180.16 g/mol
MELTING POINT:	103°C (217°F); decomposes
BOILING POINT:	Not applicable
SOLUBILITY:	Very soluble in water and acetone; soluble in ethyl alcohol and methyl alcohol

KEY FACTS

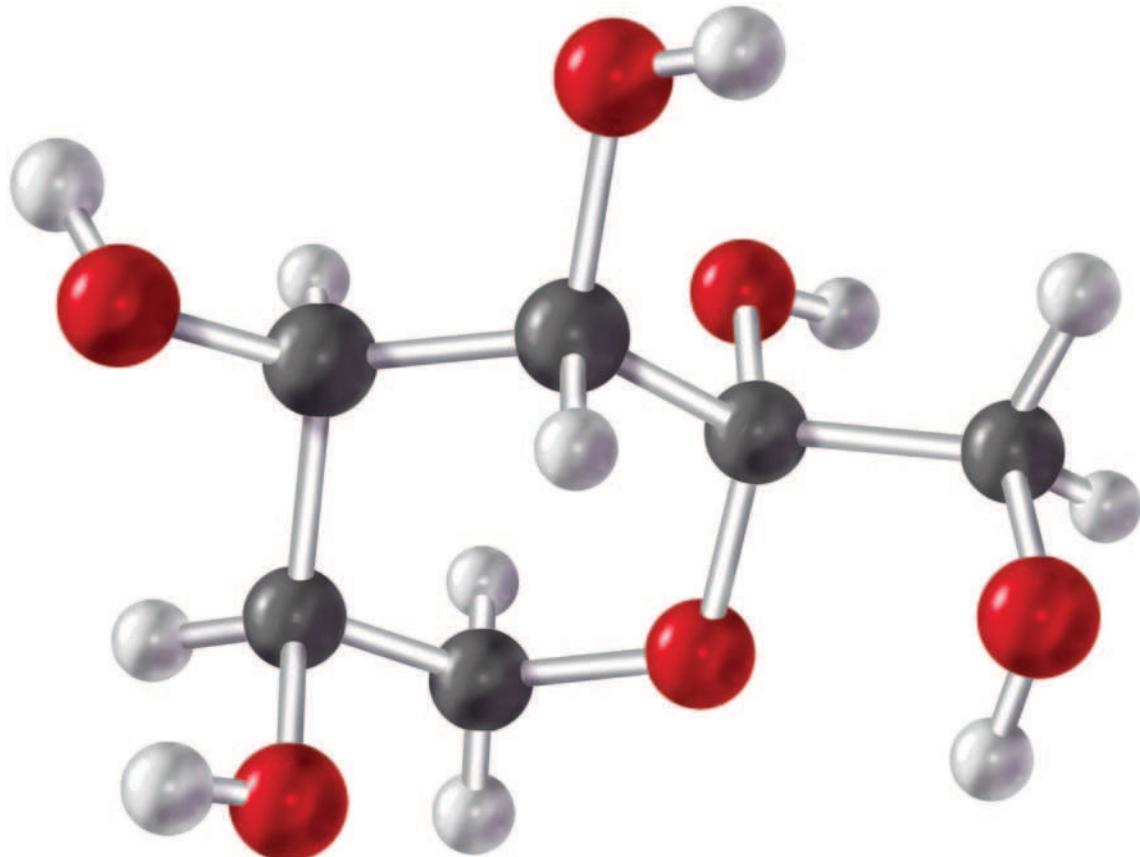
Fructose

OVERVIEW

Fructose (FROOK-tose) is a white crystalline solid found in honey and certain fruits and vegetables. It is the sweetest of the common sugars. Fructose is a carbohydrate, an organic compound in which five of the six carbon atoms are arranged in a ring to which are attached the hydrogen atoms and hydroxy (-OH) groups that make up the molecule. It is classified as a monosaccharide (“one sweet substance”), in contrast to sucrose, common table sugar, which is classified as a disaccharide (“two sweet substances”). Molecules of sucrose consist of two rings rather than the one ring found in fructose.

HOW IT IS MADE

Fructose is produced commercially by the hydrolysis of beet sugar or inulin, a polysaccharide found in the roots of a number of plants, including dahlias, Jerusalem artichokes, and chicory. Hydrolysis is the process by which a material is broken down into simpler elements by reacting it with



Fructose. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. PUBLISHERS RESOURCE GROUP

water. A polysaccharide is a carbohydrate with many simple sugar groups attached to each other. After hydrolysis of the beet sugar, or inulin, the resulting mixture is treated with lime (calcium oxide; CaO) to extract the fructose. It is then refined by removing impurities left from the preparation process.

COMMON USES AND POTENTIAL HAZARDS

Virtually the only important use of fructose is as a sweetener and preservative in a number of food products. In most cases, it is now used in the form of a substance known as high-fructose corn syrup (HFCS). HFCS was first introduced in the 1970s after scientists at the Clinton Corn Processing Company in Clinton, Iowa, developed a method of converting the sugar in corn into glucose and fructose. The Clinton

Interesting Facts

- Sucrose is a disaccharide that consists of one molecule of glucose joined to one molecule of fructose. The first step in the digestion of sucrose is the hydrolysis of the molecule, resulting in the formation of one molecule of glucose and one molecule of fructose. These compounds are the primary materials used by the human body in producing the energy needed to stay alive and grow.

process is relatively complicated. The polysaccharides in corn are first converted to glucose, and the glucose is then treated with enzymes that convert it to a thick syrup consisting of roughly half glucose and half fructose. From a nutritional standpoint, HFCS is very similar to sucrose, common table sugar, but it is less expensive to produce than sucrose and is more convenient to use in many instances.

High-fructose corn syrup has become one of the great success stories in the recent history of food processing in the United States. It has replaced sucrose in many applications, including nearly all soft drinks and fruit beverages, and in many jams and jellies, cookies, gum, baked goods, and other processed foods. Consumption of HFCS in the United States has increased from about 2 million metric tons (2.2 million short tons) in 1980 to just over 8 million metric tons (8.8 million short tons) in 2000.

When high-fructose corn syrup was first introduced, experts in nutrition did not anticipate that any health problems would be associated with the new product. After all, both fructose and glucose are naturally occurring substances that humans have been consuming for millennia. They are not so sure any more. Some evidence suggests that fructose is metabolized differently in the body than is glucose. It is not converted to energy as efficiently and may actually act more like a fat than like a sugar. Since few studies have been conducted on the ultimate chemical fate of fructose in the body, the nutritional value of HFCS is still the subject of some controversy among experts.

Words to Know

HYDROLYSIS The process by which a compound reacts with water to form two new compounds

The dangers of consuming HFCS are apparent for at least one group of people, those who lack the enzyme needed to metabolize fructose properly. Individuals with this genetic disorder may develop serious reactions if exposed to even a very small amount of fructose, reactions that include sweating, nausea, vomiting, confusion, abdominal pain, and, in extreme cases, convulsion and coma. Fortunately, this disorder is quite rare, affecting one person in about every twenty thousand individuals. For those with the disorder, however, care must be used in the kinds of sweeteners included in the diet.

FOR FURTHER INFORMATION

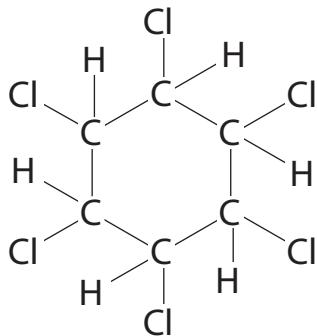
Basciano, Heather, Lisa Federico, and Khosrow Adeli. "Fructose, Insulin Resistance, and Metabolic Dyslipidemia." *Nutrition and Metabolism*. (Electronic journal) <http://www.nutritionandmetabolism.com/content/2/1/5> (accessed on October 10, 2005).

Ophardt, Charles E. "Fructose." Elmhurst College. <http://www.elmhurst.edu/~chm/vchembook/543fructose.html> (accessed on October 10, 2005).

Squires, Sally. "Sweet but Not So Innocent?" *Washington Post*. (March 11, 2003): HE01. Available online at <http://www.washingtonpost.com/ac2/wp dyn/A8003 2003Mar0?language=printer> (accessed on October 10, 2005).

"What Do We Know about Fructose and Obesity?" The International Food Information Council. <http://www.ific.org/foodinsight/2004/ja/fructosefi4o4.cfm> (accessed on October 10, 2005).

See Also Glucose; Sucrose



OTHER NAMES:

Benzene hexachloride; BHC; HCCH; HCH; TBH

FORMULA:

$C_6H_6Cl_6$

ELEMENTS:

Carbon, hydrogen, chlorine

COMPOUND TYPE:

Chlorinated aromatic hydrocarbon (organic)

STATE:

Solid

MOLECULAR WEIGHT:

290.83 g/mol

MELTING POINT:

112.5°C (234.5°F)

BOILING POINT:

323.4°C (614.1°F)

SOLUBILITY:

Insoluble in water; soluble in absolute alcohol, chloroform, and ether

**K
E
Y
F
A
C
T
S**

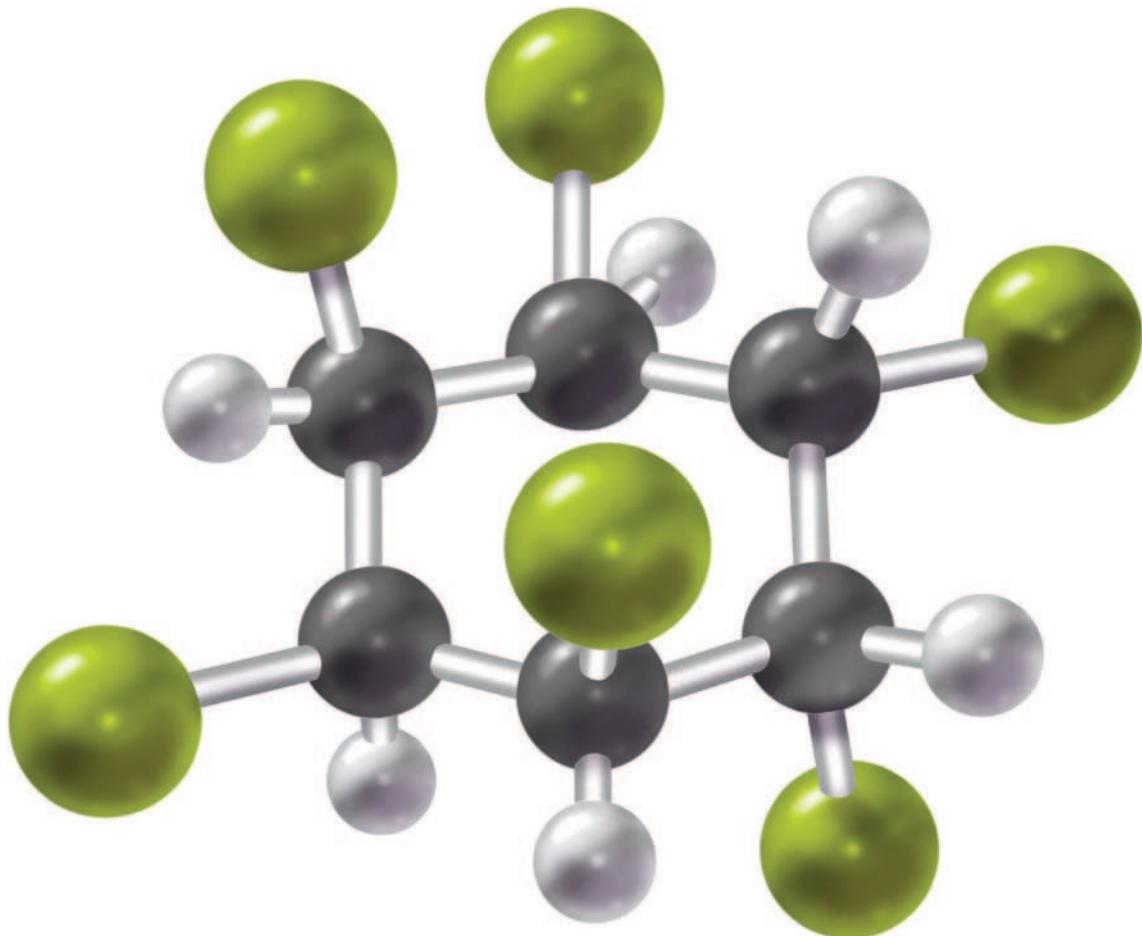
Gamma-1,2,3,4,5,6-Hexachlorocyclohexane

OVERVIEW

Gamma-1,2,3,4,5,6-hexachlorocyclohexane (GAM-uh-one two three four five six HEK-sa-KLOR-oh-SYE-klo-HEK-sane) exists in four isomeric forms. Isomers are two or more forms of a chemical compound with the same molecular formula, but different structural formulas and different chemical and physical properties. The isomer of greatest commercial interest is called the gamma (γ) isomer and is also known as lindane. It is sold commercially in more than a hundred commercial products under trade names such as 666, Africide, Agrocide, Aparasin, Arbitex, Ben-Hex, Bentox, Devoran, Entomoxan, Exagamma, Forlin, Gamaphex, Gammalin, Gam-mex, Gexane, Hexachloran, Isotox, Jacutin, Kwell, Lindafor, Lindatox, Lin-O-Sol, Lintox, Streunex, Tri-6, and Vitron.

Gamma-1,2,3,4,5,6-hexachlorocyclohexane is normally available as a white to yellowish powder or crystalline solid with a musty odor. Its color, odor, melting point, and other characteristics differ depending on the relative amount of the four isomers present in the final product. Gamma-1,2,3,4,5,

GAMMA 1,2,3,4,5,6 HEXACHLOROCYCLOHEXANE



*Gamma 1,2,3,4,5,6
Hexachlorocyclohexane.*
White atoms are hydrogen;
black atoms are carbon; and
green atoms are chlorine.

PUBLISHERS RESOURCE GROUP

6-hexachlorocyclohexane is non-flammable and stable in the presence of heat, light, strong acids, and carbon dioxide. The compound has two primary commercial uses: as a pesticide in agriculture and as a treatment for head lice, scabies, and other external parasites.

HOW IT IS MADE

Gamma-1,2,3,4,5,6-hexachlorocyclohexane is made by treating benzene (C_6H_6) with chlorine gas. In the process, each of the six hydrogen atoms in the benzene ring is replaced by a chlorine atom, resulting in the formation of a fully chlorinated benzene ring. That is, all of the original six hydrogen atoms have been replaced by chlorine atoms.

COMMON USES AND POTENTIAL HAZARDS

About 80 percent of all γ -1,2,3,4,5,6-hexachlorocyclohexane produced worldwide is used in agriculture, especially for treating soil and seeds. The wood and timber industries also use the product to protect trees from insects that attack them. In some places, γ -1,2,3,4,5,6-hexachlorocyclohexane is used as a spray to control the spread of mosquitoes. Veterinarians sometimes use the compound to treat or prevent fleas and other external parasites on animals. Gamma-1,2,3,4,5,6-hexachlorocyclohexane is also a major ingredient of products used to treat head lice, scabies, and similar pests that infest body hair.

$\text{Gamma-1,2,3,4,5,6-hexachlorocyclohexane}$ is a member of the family of compounds known as the organochlorides, organic compounds that include one or more atom of chlorine in their molecular structure. The family also includes a number of other well-known products, such as DDT, dieldrin, aldrin, endrin, dinoseb, and chlordane. These compounds kill pests by attacking and incapacitating their nervous systems. Unfortunately, they have somewhat similar effects on the nervous systems of higher animals, including humans. Many of these organochlorides have been banned in commercial applications and many nations around the world.

The status of γ -1,2,3,4,5,6-hexachlorocyclohexane is still a matter of some controversy. It is banned for medicinal uses in many countries of the world, including Bangladesh, Belize, Bolivia, Brazil, Bulgaria, Chad, Denmark, Ecuador, Egypt, Finland, Guatemala, Honduras, Hong Kong, Hungary, Indonesia, Japan, Kuwait, Mozambique, New Zealand, The Netherlands, Nicaragua, Paraguay, the Republic of Korea, Singapore, Sweden, Taiwan, and Yemen. Its use is still permitted in the United States, although it has been banned as a known carcinogen in the state of California. The product has also been banned for some or all agriculture applications in more than 50 nations.

$\text{Gamma-1,2,3,4,5,6-hexachlorocyclohexane}$ poses a health hazard to humans if ingested, inhaled, or deposited on the skin. It may cause skin irritation or rashes, nausea and vomiting, nervousness or irritability, accelerated heartbeat, convulsions or seizures, and dizziness or clumsiness. In rare cases, the ingestion of the compound has resulted in a person's death. It has also been classed as a likely carcinogen by the U.S. Environmental Protection Agency. The compound

GAMMA 1,2,3,4,5,6 HEXACHLOROCYCLOHEXANE

has been ranked in the top 10 percent among the most hazardous chemicals in ten out of eleven systems for making such rankings. It is ranked number thirty-two on the U.S. Agency for Toxic Substances and Disease Registry's list of 275 "priority" hazardous chemicals.

In spite of the apparent risk that γ -1,2,3,4,5,6-hexachlorocyclohexane poses to human health and the environment, its use is still permitted for many applications in the United States. A number of consumers' groups are currently working, however, to have the compound's use totally or partially banned.

FOR FURTHER INFORMATION

"Drug Information for Gamma Benzene Hexachloride." Drugs.com.
http://www.drugs.com/cons/Gamma_benzene_hexachloride.html
(accessed on October 10, 2005).

"Hexachlorocyclohexane (Mixed Isomers)." International Labour Organization.
http://www.ilo.org/public/english/protection/safework/cis/products/icsc/dtasht/_icsco4/icsco487.htm (accessed on October 10, 2005).

"Lindane." IPCS International Programme on Chemical Safety.
<http://www.inchem.org/documents/hsg/hsg/hsgo54.htm>
(accessed on October 10, 2005).

"Lindane Education and Research Network." National Pediculosis Association.
<http://www.headlice.org/lindane/> (accessed on October 10, 2005).

OTHER NAMES:	Gelatine
FORMULA:	Not applicable
ELEMENTS:	Carbon, hydrogen, oxygen, nitrogen, and others
COMPOUND TYPE:	Not applicable
STATE:	Solid
MOLECULAR WEIGHT:	Not applicable
MELTING POINT:	Not applicable
BOILING POINT:	Not applicable
SOLUBILITY:	Soluble in hot water and glycerol; insoluble in most organic solvents

KEY FACTS

Gelatin

OVERVIEW

Gelatin (JELL-ah-tin) is a mixture, not a compound. Mixtures differ from compounds in a number of important ways. The parts that make up a mixture are not chemically combined with each other, as they are in a compound. Also, mixtures have no definite composition, but consist of varying amounts of the substances from which they are formed. Gelatin is a mixture of water-soluble proteins with high molecular weights. It typically occurs as a brittle solid in the form of colorless or slightly yellow flakes or sheets, or in powder form, with virtually no taste or odor. It absorbs up to ten times its own weight when mixed with cold water and dissolves in hot water. When a solution of gelatin in hot water is cooled, it takes the form of a gel, a jelly-like material perhaps most commonly seen as the popular dessert called JELL-O™. Gelatin is also available in a number of other commercial forms, such as Knox Gelatin™, Puragel®, and Gelfoam®. Gelatin has been known to humans for many centuries, but it was not widely marketed until the late

1890s. Its name comes from the Latin word *gelatus*, which means “frozen.”

HOW IT IS MADE

Gelatin is made by boiling animal parts with high protein content, such as skin, ligaments, tendons, cartilage, and bones. The boiling process breaks down molecular bonds between individual collagen strands in the animal tissue. Collagen is a structural protein found in bone, cartilage, and connective tissue. The collagen formed by this process can be further disintegrated through additional boiling with either acid or alkali. Type A gelatin is produced when collagen is boiled in an acidic solution, and type B gelatin is produced by boiling collagen in an alkaline solution.

Most of the animal parts used to make gelatin come from cattle and pigs and are left over from meat and leather processing. Gelatin can also be made from fish. One of the oldest forms of gelatin is isinglass, made from the swim bladders of fish. Jewish and Muslim dietary laws prohibit believers from eating pork, so some gelatin is made without pig parts. Vegetarians and vegans do not eat any animal products, so gelatin manufacturers also make similar products using vegetable carbohydrates, such as agar and pectin. These vegetarian gelatins are not true gelatin, which is always made from animal proteins.

COMMON USES AND POTENTIAL HAZARDS

People discovered gelatin centuries ago and experimented with various uses for it. In the early 1800s, for example, gelatin was included in the food served to French soldiers as a source of dietary proteins. In the 1890s, Knox Gelatin™ was sold as a cure for dry fingernails. Manufacturers claimed that dry fingernails were caused by a lack of protein and that eating gelatin would cure the condition. No scientific evidence exists for that claim, but Knox Gelatin™ became popular among consumers nonetheless.

In 1900, the Genesee Pure Food Company began selling flavored gelatin under the name JELL-O™. In the early 1900s, the company began distributing booklets containing recipes using JELL-O™, eventually giving out more than 15 million such booklets. JELL-O™ eventually became one of the most

Interesting Facts

- Before refrigerators became common, gelatin was used to keep foods fresh and attractive. Packing a food in gelatin prevents oxygen from reacting with the food and causing spoilage.
- Synchronized swimmers sometimes use gelatin to hold their hair in place during performances.
- Aspic is a clear jelly often made with gelatin. It is a component of many elegant dishes, one of which, “Oeufs de caile en aspic et caviar” (Quail eggs in aspic with caviar), was served in first class on the doomed steam ship *Titanic* in 1912.

popular desserts in the United States and other countries. It has been used to make a variety of pleasant tasting, attractive looking desserts molded into many different shapes. Cooks have combined gelatin with water, milk, soft drinks, other liquids, whipped toppings, or mayonnaise to change its taste and texture. The product is often served with fruits or vegetables as a salad. Gelatin is also combined with marshmallows, jelly beans, jelly, yogurt, gummy candies, ice cream, and margarine to produce desserts of many textures and flavors. The product is sometimes recommended as a fat substitute because it provides volume in a diet without adding many calories. Some people include gelatin products in their diets as a way of increasing protein intake. Although plain gelatin is almost entirely protein, it actually has relatively little nutritional value.

Gelatin has many other uses, including:

- As a raw material for the manufacture of capsules and gels in the production of drugs;
- As a way of holding silver halide (silver bromide and silver iodide) crystals in place on photographic films and plates;
- In the manufacture of blocks used to determine the possible effects of various types of ammunition on human flesh;

Words to Know

ALKALI A strong base.

MIXTURE A collection of two or more elements and/or compounds with no definite composition.

PROTEIN A large, complex compound made of long chains of amino acids. Proteins have a number of essential functions in living organisms.

- As a binder that holds sand on sandpaper or to make certain types of paper products (such as playing cards) bright and shiny;
- As an additive in various types of cosmetics and skin treatments;
- In the manufacture of meshes used in the repair of wounds and in the production of artificial heart valves;
- In the production of certain types of cement;
- For the manufacture of light filters used in theatrical productions and for other specialized purposes;
- As a culturing medium for bacteria;
- As a stabilizer and thickener for certain types of foods, especially ice cream and some other dairy products;
- In the manufacture of printing inks;
- As an additive in the production of plastics and rubber products.

FOR FURTHER INFORMATION

“Gelatin.” WholeHealthMC.com.

http://www.wholehealthmd.com/refshelf/substances_view/1,1525,10151,00.html (accessed on December 22, 2005).

“A History of JELL O™ Brand.” Kraftfoods.com.

http://www.kraftfoods.com/jello/main.aspx?s=&m=jlo_history (accessed January 4, 2006).

“The Jell O Museum.” The Genesee Pure Food Co.

<http://www.jellomuseum.com/#Page1>. (accessed on December 22, 2005).

"What Exactly Is Jell O Made From?" How Stuff Works.

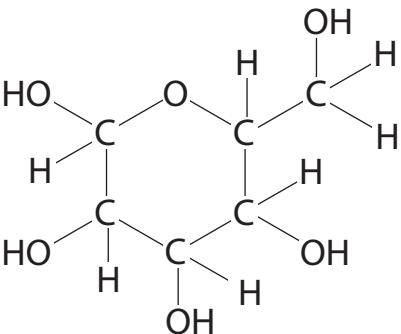
<http://home.howstuffworks.com/question557.htm> (accessed on December 22, 2005).

"What Is Gelatin." PB Leiner.

<http://www.gelatin.com/> (accessed on December 22, 2005).

Wyman, Carolyn. *JELL O: A Biography*. Fort Washington, PA: Harvest Books, 2001.

See Also Collagen



OTHER NAMES:

Dextrose; grape sugar; corn sugar; blood sugar

FORMULA:

$C_6H_{12}O_6$

ELEMENTS:

Carbon, hydrogen, oxygen

COMPOUND TYPE:

Carbohydrate (organic)

STATE:

Solid

MOLECULAR WEIGHT:

180.16 g/mol

MELTING POINT:

146°C (300°F); decomposes

BOILING POINT:

Not applicable

SOLUBILITY:

Very soluble in water; slightly soluble in alcohol; insoluble in most organic solvents

KEY FACTS

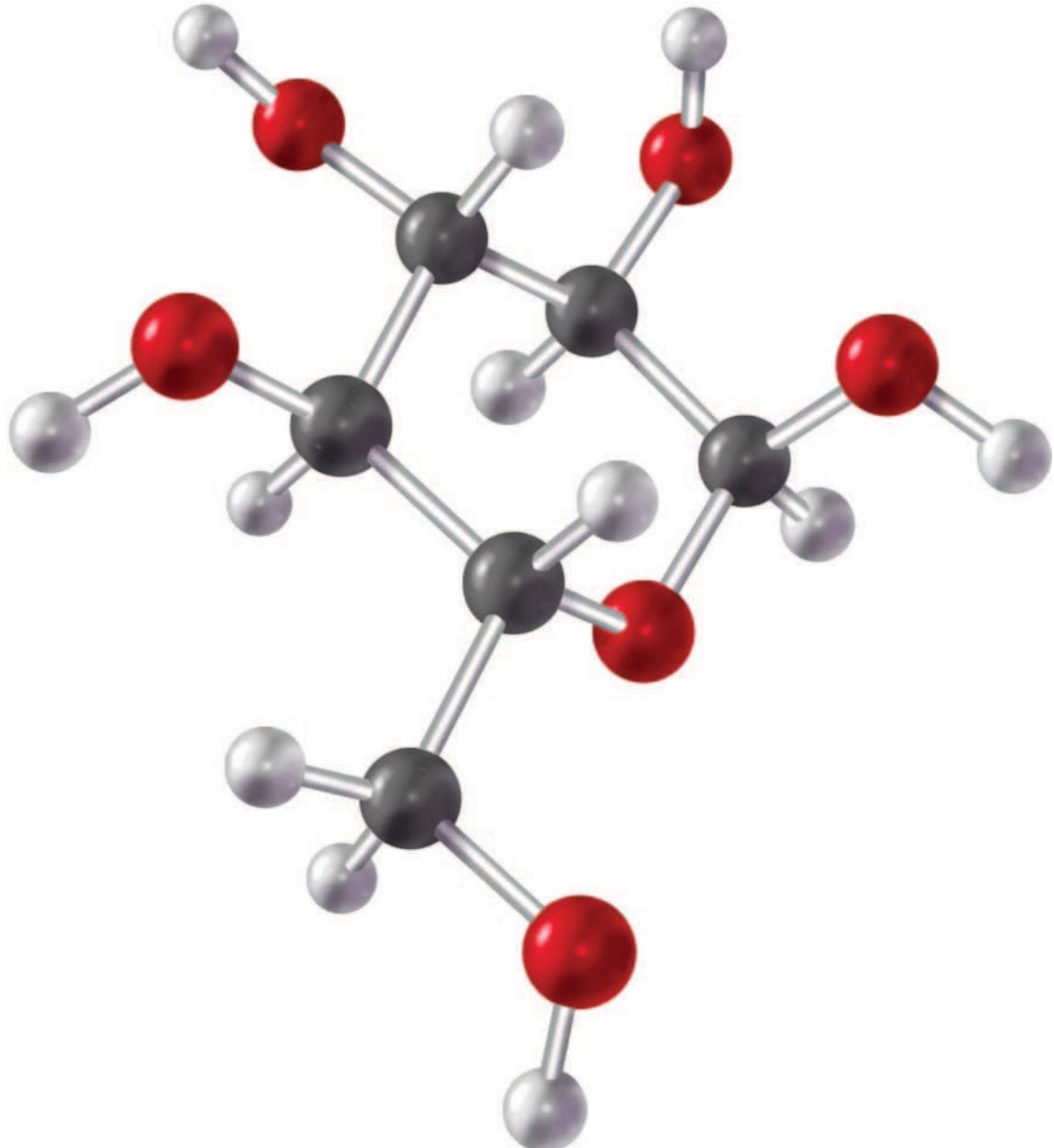
Glucose

OVERVIEW

Glucose (GLOO-kose) is a simple sugar used by plants and animals to obtain the energy they need to stay alive and to grow. It is classified chemically as a monosaccharide, a compound whose molecules consist of five- or six-membered carbon rings with a sweet flavor. Other common examples of monosaccharides are fructose and galactose. Glucose usually occurs as a colorless to white powder or crystalline substance with a sweet flavor. It consists in two isomeric forms known as the D configuration and the L configuration. Dextrose is the common name given to the D conformation of glucose.

Credit for the discovery of glucose is often given to the German chemist Andreas Sigismund Marggraf (1709–1782). In 1747, Marggraf isolated a sweet substance from raisins that he referred to as *einer Art Zucker* (a kind of sugar) that we now recognize as glucose. More than 60 years later, the German chemist Gottlieb Sigismund Constantine Kirchhof (1764–1833) showed that glucose could also be obtained from

GLUCOSE



Glucose. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. PUBLISHERS RESOURCE GROUP

the hydrolysis of starch and that starch itself was nothing other than a very large molecule (polysaccharide) composed of many repeating glucose units. The molecular structure for glucose was finally determined in the 1880s by German

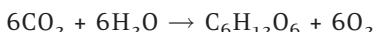
Interesting Facts

- The name glucose comes from the Greek word *gleucos* for “sweet wine.”

chemist Emil Fischer (1852–1919), part of the reason for which he was awarded the 1902 Nobel Prize in chemistry.

HOW IT IS MADE

Glucose is synthesized naturally in plants and some single-celled organisms through the process known as photosynthesis. In this process, sunlight catalyzes the reaction between carbon dioxide and water that results in the formation of a simple carbohydrate (glucose) and oxygen. The overall reaction can be summarized by a rather simple chemical equation:



However, photosynthesis actually involves a number of complex reactions that occur in two general phases, the light reactions and the dark reactions.

Glucose is produced commercially through the steam hydrolysis of cornstarch or waste products containing cellulose (a large molecule composed of glucose units) using a dilute acid catalyst. The product thus obtained is typically not very pure, but is contaminated with maltose (a disaccharide consisting of two molecules of glucose joined to each other) and dextrans (larger molecules consisting of a number of glucose units joined to each other).

COMMON USES AND POTENTIAL HAZARDS

Glucose is the primary chemical from which plants and animals derive energy. In cells, glucose is broken down in a complex series of reactions to produce energy with carbon dioxide and water as byproducts.

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

HYDROLYSIS The process by which a compound reacts with water to form two new compounds.

ISOMERS Two or more forms of a chemical compound with the same molecular formula, but different structural formulas

and different chemical and physical properties.

METABOLISM A process that includes all of the chemical reactions that occur in cells by which fats, carbohydrates, and other compounds are broken down to produce energy and the compounds needed to build new cells and tissues.

Glucose also has a number of commercial uses, nearly all of them related to the food processing business. It is used in the production of confectionary products; chewing gum; soft drinks; ice creams; jams, jellies, and fruit preparations; baby foods; baked products; and beers and ciders. A relatively small amount is used for non-food purposes, primarily in the production of other organic chemicals, such as citric acid, the amino acid lysine, insulin, and a variety of antibiotics.

The most important health problem associated with glucose is diabetes. Diabetes is a medical condition that develops when the body either does not produce adequate amounts of insulin or cannot use that compound properly. Insulin is a hormone that controls the metabolism of glucose in the body. If glucose is not metabolized properly, a person's body acts as if it is "starving." Symptoms of diabetes include excessive hunger, weight loss, and exhaustion. If left untreated, the condition can result in coma and death. Diabetics must have an artificial source of insulin (usually from injections) and watch their diets to keep these symptoms under control.

FOR FURTHER INFORMATION

"All about Diabetes." American Diabetes Association.
<http://www.diabetes.org/about-diabetes.jsp> (accessed on October 10, 2005).

“Carbohydrates.” Kimball’s Biology Pages.

<http://users.rcn.com/jkimball.ma.ultranet/BiologyPages/C/Carbohydrates.html> (accessed on October 10, 2005).

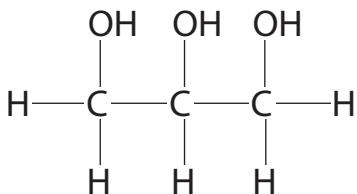
“Dextrose, Anhydrous.” J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/Do835.htm> (accessed on October 10, 2005).

“Glucose.” Department of Chemistry, Imperial College London.

http://www.ch.ic.ac.uk/vchemlib/mim/bristol/glucose/glucose_text.htm (accessed on October 10, 2005).

See Also Fructose; Sucrose



OTHER NAMES:

Glycerin; glycerine;
glycyl alcohol

FORMULA:

$\text{CH}_2\text{OHCHOHCH}_2\text{OH}$

ELEMENTS:

Carbon, hydrogen,
oxygen

COMPOUND TYPE:

Trihydric alcohol
(organic)

STATE:

Liquid

MOLECULAR WEIGHT:

92.09 g/mol

MELTING POINT:

18.1°C (64.6°F)

BOILING POINT:

290°C (554°F)

SOLUBILITY:

Miscible with water
and alcohol; insoluble
in ether, benzene, and
chloroform

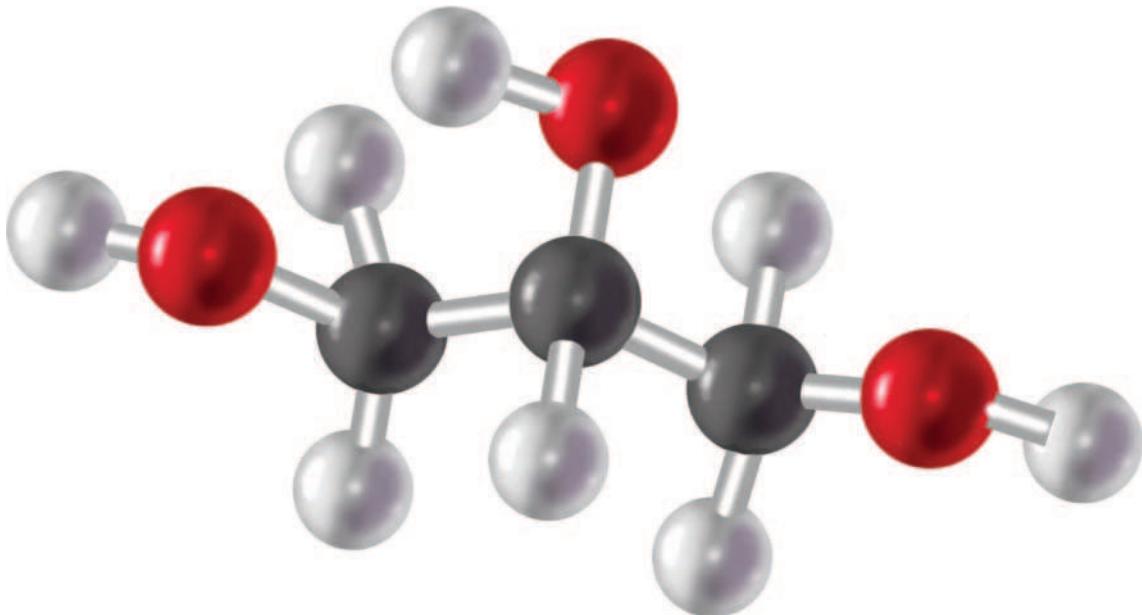
KEY FACTS

OVERVIEW

Glycerol (GLIH-ser-ol) is a clear, colorless, odorless, sweet-tasting syrupy liquid. It is a trihydric alcohol, meaning that its molecules contain three hydroxyl (-OH) groups. Glycerol occurs naturally in all animal and plant cells. Glycerol was discovered in 1779 by the German chemist Karl Wilhelm Scheele (1742–1786) and named by the French chemist Michel Eugéne Chevreul (1786–1889) because of its sweet taste (*glycos* means “sweet” in Greek). In 1836 the French chemist Théophile-Jules Pelouze (1807–1867) determined the molecular formula for glycerol, and three decades later, in 1872, the compound was first synthesized (created in a laboratory) by the French chemist Charles Friedel (1832–1899). About 250 million kilograms (500 million pounds) of glycerol are produced in the United States each year, the majority of which goes to the production of food and personal care products.

HOW IT IS MADE

The traditional method of making glycerol is by the saponification of fats. Fats are the esters of glycerol and



Glycerol. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. PUBLISHERS RESOURCE GROUP

one or more alcohols. When a fat is hydrolyzed in the presence of a catalyst, it is converted into the glycerol and alcohols from which it was originally made. This type of hydrolysis is called saponification because it is the usual method by which soaps (the Latin word for “soap” is *sapon*). Glycerol can be obtained, then, as a byproduct in the manufacture of soaps.

A number of synthetic methods for making glycerol are also available. Most of these procedures begin with propylene (propene; $\text{CH}_2=\text{CHCH}_3$) and chlorine and involve a series of steps that convert the three-carbon propylene to the three-carbon glycerol. Glycerol can also be obtained by treating simple sugars with hydrogen over a nickel catalyst. In the United States, about 80 percent of all glycerol produced is obtained by saponification methods, and the remaining 20 percent is made by synthetic methods.

COMMON USES AND POTENTIAL HAZARDS

The most common use for glycerol in the United States is in food products, where it acts as a sweetener and as a thickener in many foods. For example, it is added to ice cream to improve texture and to candy products and baked

Interesting Facts

- As winter approaches, many insects begin to produce glycerol to replace the water in their body tissues. Glycerol acts as an antifreeze to prevent the insects from freezing during the coldest part of the year.

goods to increase the sweetness of the product. It is also used to make the flexible coatings on cheeses, sausages, and other meat products. About one-quarter of all glycerol made in the United States is used in the food products industry.

Nearly the same amount of glycerol is used in the preparation of personal care products, such as skin, hair, and soap products (23 percent) and in oral hygiene products, such as toothpastes and mouthwashes (17 percent). Some of the products that include glycerol are moisturizers, detergents, soaps, hair coloring agents, mascara, nail polish, nail polish removers, perfumes, body lotions, hair sprays, shaving creams, lipsticks, cough medicines, shampoos, and hair conditioners. Glycerol is also used as a humectant in tobaccos. A humectant is a material that helps a product conserve moisture and prevent it from drying out.

Other uses for glycerol include:

- In the manufacturer of explosives;
- In the production of a variety of plastics and polymers, such as polyether polyols, urethanes, and alkyd resins;
- As a lubricant in pumps, bearings, gaskets, and other mechanical systems;
- In the manufacture of ink rolls, inks, and rubber stamps;
- As an emulsifying agent, a material that helps two liquids that are not soluble in each other to stay mixed;
- As an antifreeze; and
- In a number of medical applications, such as the treatment of glaucoma and stroke.

Words to Know

ESTER An organic compound formed in the reaction between an organic acid and an alcohol.

HYDROLYSIS The process by which a compound reacts with water to form two new compounds.

MISCIBLE able to be mixed; especially applies to the mixing of one liquid with another.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

Glycerol poses some safety problems because it is combustible and explosive under certain conditions. It presents no health hazards under normal circumstances of use, however.

FOR FURTHER INFORMATION

“Glycerol.” International Chemical Safety Cards.

<http://www.healthycommunications.com/msdsglycerin1.html>
(accessed on October 10, 2005).

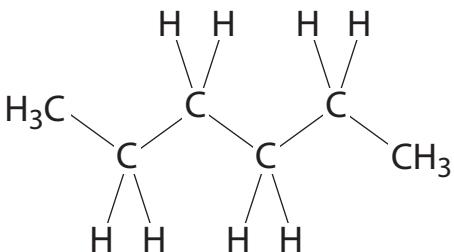
“Glycerol.” PDRhealth.

http://www.pdrhealth.com/drug_info/nmdrugprofiles/nutsupdrugs/gly_0304.shtml (accessed on October 10, 2005).

Legwold, Gary. “Hydration Breakthrough.” *Bicycling* (July 1994): 72–73.

“Material Safety Data Sheet: Glycerine.” Department of Chemistry, Iowa State University.

<http://avogadro.chem.iastate.edu/MSDS/glycerine.htm>
(accessed on October 10, 2005).



OTHER NAMES:	n-hexane
FORMULA:	C_6H_{14}
ELEMENTS:	Carbon, hydrogen
COMPOUND TYPE:	Alkane; saturated hydrocarbon (organic)
STATE:	Liquid
MOLECULAR WEIGHT:	86.18 g/mol
MELTING POINT:	-95.35°C (-139.6°F)
BOILING POINT:	68.73°C (155.7°F)
SOLUBILITY:	Insoluble in water; very soluble in ethyl alcohol; soluble in ether and chloroform

KEY FACTS

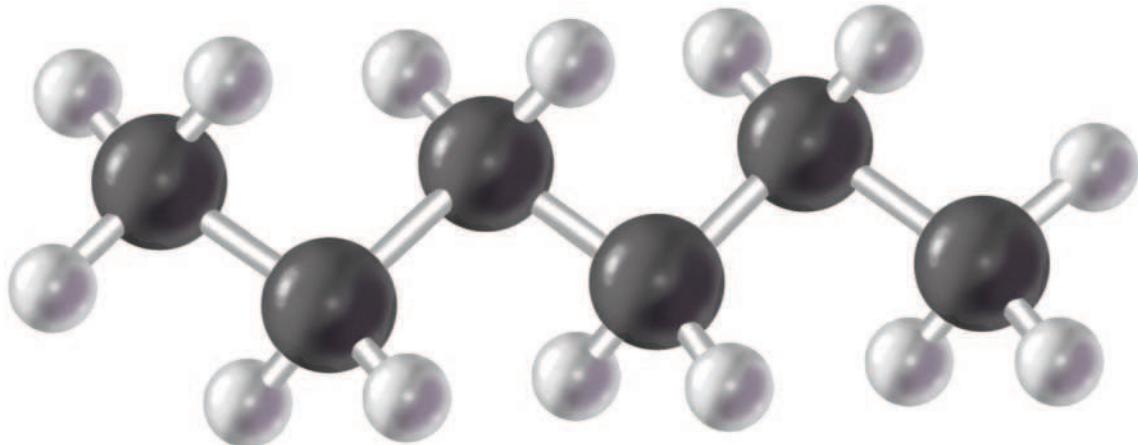
Hexane

OVERVIEW

Hexane (HEX-ane) is a colorless flammable liquid with a faint petroleum-like odor. Chemically it is classified as a saturated hydrocarbon, which means that its molecules contain only carbon and hydrogen atoms joined only by single bonds. Saturated hydrocarbons are also known as alkanes. By far its most important use is as a solvent in a variety of industrial operations.

HOW IT IS MADE

Hexane is extracted from petroleum. Petroleum is a complex mixture of solid, liquid, and gaseous hydrocarbons that has virtually no use itself. However, the fractional distillation of petroleum produces hundreds of individual compounds, each of which has its own important commercial and industrial applications. Fractional distillation is the process by which petroleum is heated in tall towers. The components of petroleum boil off at different temperatures, rise to



Hexane. White atoms are hydrogen and black atoms are carbon. PUBLISHERS RESOURCE GROUP

different heights in the tower, and are condensed at different levels above the base of the tower. As a liquid with a low boiling point, hexane boils off and rises to upper levels of the tower, where it condenses and is removed in a portion of the petroleum known as petroleum ether or ligroin. Hexane can then be separated from other constituents of petroleum ether by a second distillation, in which each component boils off and is condensed at its own distinctive boiling point.

COMMON USES AND POTENTIAL HAZARDS

By far the most important use of hexane is in solvents used for a variety of purposes, such as the extraction of oils from seeds and vegetables; as a degreaser and cleaning agent for printing equipment; as a solvent in glues such as rubber cement; as an ingredient in inks and varnishes; in the shoe and leather manufacturing industry; and in the roofing industry.

Hexane poses both safety and health risks for humans and other animals. The liquid vaporizes easily and the vapors formed ignite easily and may even explode under the proper conditions. The primary health hazard related to hexane occurs by breathing in the compound. When inhaled, it can cause numbness in the hands and feet, weakness in the feet and lower legs, paralysis of the arms and legs, muscle wasting,

Interesting Facts

- Hexane to which a red or blue dye has been added is sometimes used to make thermometers used for measuring low temperatures.

damage to nerves, nausea and vomiting, jaundice, skin rashes, irritation of the eyes and throat, blurred vision, mental confusion, and coma. There is no evidence, however, that hexane is carcinogenic.

The health hazards posed by hexane are of concern in only two circumstances: among workers who handle the liquid on a regular basis; and among people who deliberately inhale the compound as part of a solvent for the purpose of getting “high.” In both cases, a person is exposed to much higher concentrations of hexane than one would encounter in commercial or industrial products. The practice of glue-sniffing, which has become popular among some teenagers, can result in serious health problems, including dizziness, increased heart rate, high blood pressure, disorientation, confusion, and occasionally violent impulses or suicide attempts.

Words to Know

CARCINOGEN A substance that causes cancer in humans or other animals.

sing the vapors produced at different temperatures.

DISTILLATION A process of separating two or more substances by boiling the mixture of which they are composed and conden-

SOLVENT A liquid in which another substance is dissolved, to form a solution.

FOR FURTHER INFORMATION

"Hexane." U.S. Environmental Protection Agency Technology Transfer Network, Air Toxics Website.

<http://www.epa.gov/ttn/atw/hlthef/hexane.html> (accessed on October 12, 2005).

Menhard, Francha Roffe. *The Facts about Inhalants*. New York: Benchmark Books, 2004.

"ToxFAQs for n hexane." Agency for Toxic Substances and Disease Registry.

<http://www.atsdr.cdc.gov/tfacts113.html> (accessed on October 12, 2005).

**OTHER NAMES:**

Anhydrous
hydrochloric acid

FORMULA:

HCl

ELEMENTS:

Hydrogen; chlorine

COMPOUND TYPE:

Inorganic acid

STATE:

Gas

MOLECULAR WEIGHT:

36.46 g/mol

MELTING POINT:

-114.17°C (-173.51°F)

BOILING POINT:

-85°C (-121°F)

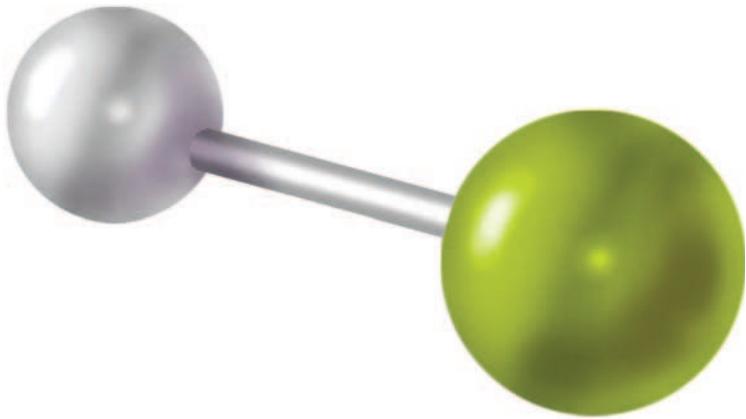
SOLUBILITY:

Very soluble in
water; soluble in
alcohol and ether

K E Y F A C T S**OVERVIEW**

Hydrogen chloride (HY-druh-jin KLOR-ide) is a colorless gas with a strong, suffocating odor. The gas is not flammable, but is corrosive, that is, capable of attacking and reacting with a large variety of other compounds and elements. Hydrogen chloride is most commonly available as an aqueous solution known as hydrochloric acid. It is one of the most important industrial chemicals in the world. In 2004, just over 5 million metric tons (5.5 million short tons) of hydrogen chloride were produced in the United States, making it the eighteenth most important chemical in the nation for that year.

Hydrogen chloride has probably been known as far back as the eighth century, when the Arabian chemist Jabir ibn Hayyan (c. 721-c. 815; also known by his Latinized name of Geber) described the production of a gas from common table salt (sodium chloride; NaCl) and sulfuric acid (H_2SO_4). The compound was mentioned in the writings of a number of alchemists during the Middle Ages and was probably first



Hydrogen chloride. White atom is hydrogen and green atom is chlorine. PUBLISHERS RESOURCE GROUP

produced in a reasonably pure form by the German chemist Johann Rudolf Glauber (1604-1670) in about 1625. The first modern chemist to prepare hydrogen chloride and describe its properties was the English chemist Joseph Priestley (1733-1804) in 1772. Forty years later, in 1818, the English chemist and physicist Humphry Davy (1778-1829) showed that the compound consisted of hydrogen and chlorine, giving it the correct formula of HCl.

Commercial production of hydrogen chloride had its beginning in Great Britain in 1823. The method of production most popular there and, later, throughout Europe was one originally developed by the French chemist Nicholas Leblanc (1742-1806) in 1783. Leblanc had invented the process as a method for producing sodium hydroxide and sodium carbonate, two very important industrial chemicals. Hydrogen chloride was produced as a byproduct of the Leblanc process, a byproduct for which there was at first no use. The gas was simply allowed to escape into the air. The suffocating and hazardous release of hydrogen chloride prompted governments to pass legislation requiring some other means of disposal for the gas. In England, that law was called the Alkali Act and was adopted by the parliament in 1863. Unable to release hydrogen chloride into the air, manufacturers began dissolving it in water and producing hydrochloric acid. Before long, a number of important commercial and industrial uses for the acid itself were discovered. The “useless” byproduct of the Leblanc process soon became as important as the primary products of the process, sodium hydroxide and sodium carbonate.

Interesting Facts

- Hydrogen chloride was studied by the famous alchemist Basil Valentine (c. 1394–?), who gave it the name *spiritus salis* (“the spirit of salt”) by which it was known to most alchemists.
- Hydrochloric acid has traditionally been known as muriatic acid, a name that is still sometimes used by workers in fields in which it is used.

HOW IT IS MADE

Hydrogen chloride is still sometimes made today by the traditional process of reacting sodium chloride (NaCl) with a sulfate, such as sulfuric acid or iron(II) sulfate (FeSO_4). However, more than 90 percent of the hydrogen chloride produced throughout the world today comes as the byproduct of the chlorination of organic compounds. Chlorination is the process by which chlorine gas reacts with an organic compound, usually replacing some of the hydrogen present in the compound. Since a large number of important chlorinated organic compounds are produced each year, large amounts of hydrogen chloride gas are produced as a byproduct. That gas is simply removed from the reaction and stored in cylinders for future use. Other methods of producing hydrogen chloride include the direct synthesis of hydrogen gas and chlorine gas (producing a very pure product) and the reaction of sodium chloride, sulfur dioxide, oxygen, and water with each other at high temperatures (the Hargreaves process).

COMMON USES AND POTENTIAL HAZARDS

Hydrogen chloride and hydrochloric acid have some uses in common, and some that are different from each other. In both dry and liquid form, the largest single use of hydrogen chloride is in the synthesis of organic and inorganic chlorides. A large number of compounds important in commerce and industry contain chlorine, including most pesticides, many pharmaceuticals, and a number of polymeric products.

Words to Know

ALCHEMY An ancient field of study from which the modern science of chemistry evolved.

ANHYDROUS Without water or moisture.

AQUEOUS A solution is one that consists of some material dissolved in water.

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

Hydrochloric acid is also used widely in the processing of metallic ores and the pickling of metals. Pickling is the process by which a metal is cleaned, usually with an acid, to remove rust and other impurities that have collected on the metal. Some additional uses of hydrogen chloride and hydrochloric acid include the following:

- In the brining of foods and other materials. Brining is the process by which a material is soaked in a salt solution, usually in order to preserve the material;
- In the treatment of swimming pool water;
- As a catalyst in industrial chemical reactions;
- In the manufacture of semiconductors and other electronic components;
- To maintain the proper acidity in oil wells (to keep oil flowing smoothly);
- For the etching of concrete surfaces;
- In the production of aluminum, titanium, and a number of other important metals.

Both hydrogen chloride and hydrochloric acid pose serious health risks to humans and other animals. The gas is an irritant to the eyes and respiratory system, causing coughing, choking, and tearing, as well as more serious damage to tissues. Hydrochloric acid can burn the skin and mucous membranes. Exposure of only five parts per million of the gas can produce noticeable symptoms of distress, and exposure of more than 2,000 parts per million can be fatal. If hydrochloric

acid gets into the eyes, blindness may result. Since hydrochloric acid is present in many household products, users should exercise great care when working with such materials.

FOR FURTHER INFORMATION

“Hydrochloric Acid.” National Safety Council.

<http://www.nsc.org/library/chemical/Hydrochl.htm> (accessed on October 12, 2005).

“Hydrogen Chloride.” Agency for Toxic Substances and Disease Registry.

<http://www.atsdr.cdc.gov/tfacts173.pdf> (accessed on October 12, 2005).

“Hydrogen Chloride.”

<http://www.ucc.ie/ucc/depts/chem/dolchem/html/comp/hcl.html> (accessed on October 12, 2005).

“Hydrogen Chloride, HCl.” Defense Service Center.

http://www.cfc.com/gaslink/pure/hydrogen_chloride.htm (accessed on October 12, 2005).

See Also Sodium Chloride



OTHER NAMES:	
Hydrogen dioxide;	
hydroperoxide;	
peroxide	
FORMULA:	
H_2O_2	
ELEMENTS:	
Hydrogen; oxygen	
COMPOUND TYPE:	
Oxide (inorganic)	
STATE:	
Liquid	
MOLECULAR WEIGHT:	
34.02 g/mol	
MELTING POINT:	
-0.43°C (-31°F)	
BOILING POINT:	
150.2°C (302.4°F)	
SOLUBILITY:	
Very soluble in water; soluble in ether	

KEY FACTS

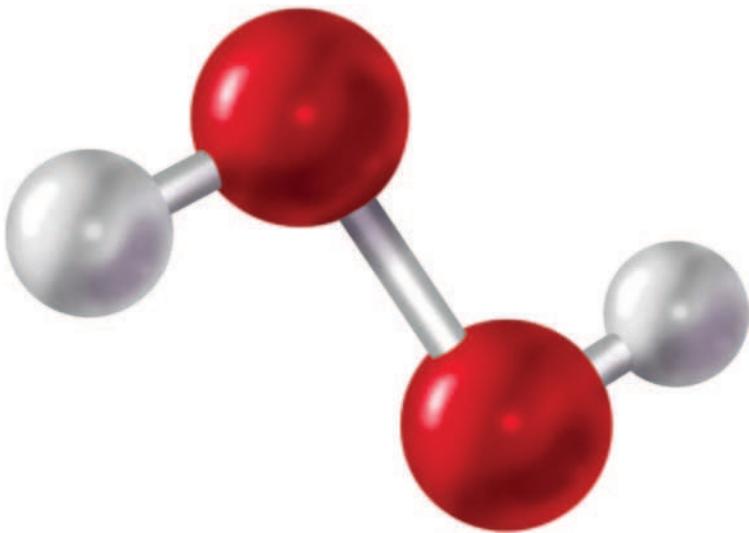
Hydrogen Peroxide

OVERVIEW

Hydrogen peroxide (HY-druh-jin per-OK-side) is a clear, colorless, somewhat unstable liquid with a bitter taste. When absolutely pure, the compound is quite stable. Even small amounts of impurities (such as iron or copper), however, act as catalysts that increase its tendency to decompose, sometimes violently, into water and nascent oxygen (O). To prevent decomposition, small amounts of inhibitors, such as acetanilide or sodium stannate are added to pure hydrogen peroxide and hydrogen peroxide solutions.

Hydrogen peroxide was discovered in 1818 by French chemist Louis Jacques Thénard (1777-1857). It was first used commercially in the 1800s, primarily to bleach hats. Today, industrial processes make about 500 million kilograms (1 billion pounds) of hydrogen peroxide annually for use in a wide variety of applications ranging from whitening of teeth to propelling rockets.

Hydrogen peroxide. Red atoms are oxygen and white atoms are hydrogen. PUBLISHERS RESOURCE GROUP



HOW IT IS MADE

Hydrogen peroxide occurs in very small amounts in nature. It is formed when atmospheric oxygen reacts with water to form H_2O_2 . Hydrogen peroxide is also present in plant and animal cells as the byproduct of metabolic reactions that occur in those cells.

The large amounts of hydrogen peroxide used in industry are prepared in a complex series of reactions that begins with any one of a family of compounds known as the alkyl anthrahydroquinones, such as ethyl anthrahydroquinone. The anthrahydroquinones are three-ring compounds that can be converted back and forth between two or more similar structures. During the conversion from one structure to another, hydrogen peroxide is produced as a byproduct. The anthraquinone is continuously regenerated during the production of hydrogen peroxide, making the process very efficient.

Other methods for the preparation of hydrogen peroxide are also available. For example, the electrolysis of sulfuric acid results in the formation of a related compound, peroxy-sulfuric acid (H_2SO_5), which then reacts with water to form hydrogen peroxide. A third method of preparation involves the heating of isopropyl alcohol [2-propanol; $(\text{CH}_3)_2\text{CHOH}$] at high temperature and pressure, resulting in the formation of hydrogen peroxide as one product of the reaction.

Interesting Facts

- Hydrogen peroxide is sold in concentrations ranging from 3 percent (for home use) to 90 percent (for industrial applications).
- Scientists have discovered hydrogen peroxide in the atmosphere of Mars.

COMMON USES AND POTENTIAL HAZARDS

Most of hydrogen peroxide's applications depend on the fact that it tends to break down, releasing a single atom of nascent oxygen (O):



The term nascent oxygen refers to a single atom of oxygen, a structure that is chemically very active. Nascent oxygen tends to be a very strong oxidizing agent. For example, the use of hydrogen peroxide with which most people are probably familiar is as an antiseptic, a substance used to kill germs. Hydrogen peroxide achieves this result because the nascent oxygen it releases destroys bacteria, fungi, and other microorganisms that cause disease.

The most important industrial application of hydrogen peroxide—its use in the pulp and paper industry—also depends on its oxidizing properties. In this case, it is used to bleach the materials of which paper is made, converting colored compounds to colorless compounds. About 55 percent of all hydrogen peroxide made in the United States is used for this purpose. Another nine percent is used in the bleaching of other materials, such as textiles, furs, feathers, and hair. Another important application of hydrogen peroxide is in water and sewage treatment plants, where its antibacterial action destroys disease-causing organisms in the water. Some additional uses of hydrogen peroxide include:

- In bakeries to condition dough and make it easier to work with;
- For cleaning metals;
- As a rocket propellant;

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

INHIBITOR A substance added to another substance to prevent or slow down an unwanted reaction.

METABOLISM A process that includes all of the chemical reactions that occur in cells

by which fats, carbohydrates, and other compounds are broken down to produce energy and the compounds needed to build new cells and tissues.

OXIDATION A chemical reaction in which oxygen reacts with some other substance or, alternatively, in which some substances loses electrons to another substance, the oxidizing agent.

- In the preparation of other organic and inorganic compounds;
- As a neutralizing agent in the production of wines; and
- As a disinfectant in the treatment of seeds for agricultural purposes.

The hydrogen peroxide solutions with which people come into contact at home pose little or no health hazard because the concentration of the compound is very low, usually about 3 percent. Prolonged use of hydrogen peroxide may cause burns on the skin, however, and the more concentrated solutions used in industry present more serious hazards. They can be toxic if ingested and are explosive if not stored properly.

FOR FURTHER INFORMATION

“Hydrogen Peroxide (>60% Solution in Water).” International Labour Organization.
http://www.ilo.org/public/english/protection/safework/cis/products/icsc/dtasht/_icsco1/icsco164.htm
(accessed on October 12, 2005).

“Introduction to Hydrogen Peroxide.” U.S. Peroxide.
<http://www.h2o2.com/intro/overview.html>
(accessed on October 12, 2005).

“A Prescription for Death?” CBSNews.com
<http://www.cbsnews.com/stories/2005/01/12/6oII/main666489.shtml> (accessed on October 12, 2005).

OTHER NAMES:	Ferrous oxide; iron monoxide
FORMULA:	FeO
ELEMENTS:	Iron, oxygen
COMPOUND TYPE:	Metal oxide
STATE:	Solid
MOLECULAR WEIGHT:	71.84 g/mol
MELTING POINT:	1,377°C (2,151°F)
BOILING POINT:	Not applicable
SOLUBILITY:	Insoluble in water and all ordinary organic solvents; soluble in acids

K E Y F A C T S

Iron(II) Oxide

OVERVIEW

Iron(II) oxide (EYE-urn two OK-side) is a black, dense powder that occurs in nature as the mineral wustite. It reacts readily with oxygen in the air to form iron(III) oxide and with carbon dioxide to form iron(II) carbonate (FeCO_3).

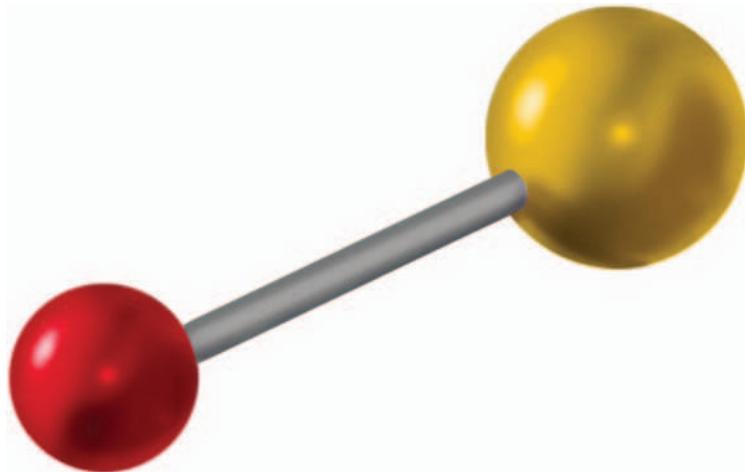
HOW IT IS MADE

Iron(II) oxide occurs in nature as the result of the incomplete oxidation of iron metal. It can be prepared synthetically by heating iron(II) oxalate (FeC_2O_4), although the product of this reaction is contaminated with another oxide of iron, triiron tetroxide (Fe_3O_4).

COMMON USES AND POTENTIAL HAZARDS

Iron(II) oxide has three primary uses. First, it has long been used as a dye or pigment in pottery, glazes, and glasses, especially in the green, heat-absorbing glass used in buildings, automobiles, and other applications. The compound is

Iron(II) oxide. Red atom is oxygen and orange atom is iron. Gray stick is a double bond. PUBLISHERS RESOURCE GROUP



also used as a raw material in the production of steel. Finally, iron(II) oxide is used as a catalyst in a number of industrial and chemical operations.

Inhalation of iron(II) oxide fumes or dust is considered a hazard and can cause throat and nasal irritation. High levels of exposure may lead to a condition known as metal fume fever, a workplace exposure illness that causes flu-like symptoms. Continued exposure at high levels can have more serious effects, including a disease known as siderosis, an inflammation of the lungs that is accompanied by pneumonia-like symptoms.

Interesting Facts

- Both iron(II) oxide and its cousin, iron(III) oxide are abundant in rocks on the surface of Mars. The former accounts for the dark, nearly jet black, color

of some rocks, while the latter is responsible for the predominant red color of the planet's surface.

Words to Know

OXIDATION A chemical reaction in which oxygen reacts with some other substance or, alternatively, in which some substances

loses electrons to another substance, the oxidizing agent.

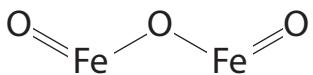
FOR FURTHER INFORMATION

Cornell, Rochelle M., and Udo Schwertmann. *The Iron Oxides; Structure, Properties, Reactions, and Uses*. Second edition. New York: Wiley VCH, 2003.

"Ferrous Oxide." International Programme for Chemical Safety.
<http://www.inchem.org/documents/icsc/icsc/eicso793.htm>
(accessed on October 12, 2005).

"Material Safety Data Sheet." ESPI Metals.
<http://www.espimetals.com/msds's/ironoxidefeo.pdf>
(accessed on October 12, 2005).

See Also Iron(III) Oxide



OTHER NAMES:
Ferric oxide; red iron oxide; red iron trioxide
FORMULA:
Fe_2O_3
ELEMENTS:
Iron; oxygen
COMPOUND TYPE:
Metallic oxide
STATE:
Solid
MOLECULAR WEIGHT:
159.69 g/mol
MELTING POINT:
1,565°C (2,849°F)
BOILING POINT:
Not applicable
SOLUBILITY:
Insoluble in water and all conventional organic solvents; soluble in acids

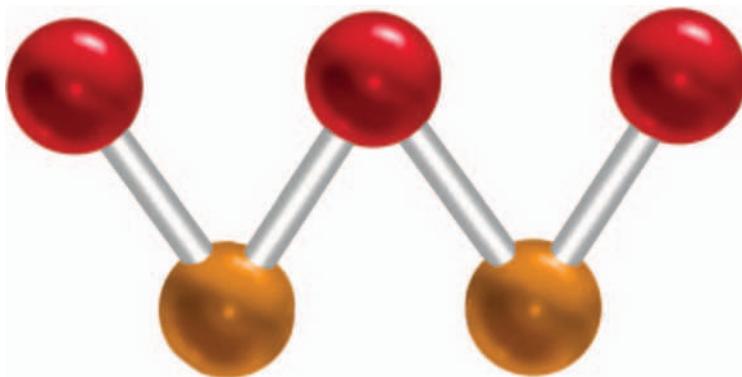
KEY FACTS

OVERVIEW

Iron(III) oxide (EYE-urn three OK-side) is a dense, reddish-brown, crystalline compound that usually occurs as lumps or a powder. It occurs in nature as the mineral hematite and is a component of the rust that forms on objects made out of iron that are exposed to the air. Rust itself is actually a complex mixture of iron oxides and hydroxides, including Fe_2O_3 , FeO , Fe_3O_4 and FeO(OH) . Hematite may range in color from black and silver gray to reddish brown and red depending on the type and amount of impurities present with iron(III) oxide. Iron(III) is also ferromagnetic. Ferromagnetism refers to the ability of a substance to become highly magnetic and then retain its magnetism.

HOW IT IS MADE

Hematite forms naturally when iron-containing rocks and minerals react with oxygen in the air to form iron(III) oxide. The oxide can be made synthetically by a variety of



Iron(III) oxide. Red atoms are oxygen and orange atoms are iron. PUBLISHERS RESOURCE GROUP

procedures. In the most popular method, iron(II) sulfate is reacted with sodium hydroxide (NaOH) to produce iron(II) hydroxide [$\text{Fe}(\text{OH})_2$]. The iron(II) hydroxide is then allowed to react with oxygen in the air, forming iron(III) oxide. The compound can also be produced by heating iron(II) sulfate, hydrated iron(II) oxide ($\text{FeO}(\text{OH})$), or iron(III) oxalate [$\text{Fe}_2(\text{C}_2\text{O}_3)_3$].

COMMON USES AND POTENTIAL HAZARDS

Iron(III) oxides has been associated with the manufacture of iron and steel for much of human history. The Iron Age, which began in Egypt around 4000 BCE was the period in human history when iron was used for tools and weapons. The general approach to refining iron metal from iron ores, such as hematite, was to heat the ore in the presence of carbon. Carbon removes oxygen from the ore, leaving the free metal behind. By the first century BCE in China, the first known blast furnaces were in use. In a blast furnace, iron(III) oxide is reduced with carbon by using a blast of air and heat. The oxygen from the air reacts with carbon to give carbon monoxide, which then reacts with iron(III) oxide to produce liquid iron metal and carbon dioxide.

In the eighteenth century, the blast furnace process was further developed so that iron could be made commercially. This process can be traced to the region around

Interesting Facts

- The name hematite is derived from the Greek word for blood.
- When hematite is made into an ornament, it is sometimes called black diamond.
- NASA's Mars rover *Opportunity* found small particles thought to be mostly hematite on the planet's surface. Scientists think they formed billions of

years ago when Mars had water on its surface.

- Paleolithic humans in Swaziland, who lived more than 40,000 years ago, mined hematite in the oldest known mine in the archaeological record called the Lion Cave. It is thought that they mined the hematite to produce the red pigment known as ochre.

Coalbrookdale in Shropshire, England, around the year 1773 and is said to have been a factor in initiating the Industrial Revolution. The blast furnace method is still one of the primary methods by which iron metal is refined from iron ores.

Iron(III) oxide is also one of the oldest known pigments and has been used for that purpose in every major civilization. Some of the best known pigments made from iron(III) oxide have been Indian red, terra Pozzuoli, and Venetian red and have been used to color ceramic glazes and paints. Depending on the exact formulation used, iron(III) oxide produces colors ranging from yellow to orange to red. For example, the hydrated oxide produces a pigment ranging from yellow to brown. Iron(III) oxide pigments have been used as pigment for rubber, paper, linoleum, glass, and many types of paints, including specialty paints used on metalwork and ship hulls.

Some of the other commercial and industrial applications of iron(III) oxide include:

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

MORDANT A substance used in dyeing and printing that reacts chemically with both a dye and the material being dyed to help hold the dye permanently to the material.

- As a catalyst for many industrial and chemical operations;
- As a component of thermite, a mixture of iron(III) oxide and aluminum powder which, when ignited, produces very hot temperatures. Thermite bombs are used in welding;
- In computer hard disks, audio cassette tapes, video cassette tapes, and computer floppy disks for magnetic storage of data;
- As an abrasive and polish for use with brass, steel, gems, and other hard objects;
- As a mordant in the dyeing of cloth;
- As a feed additive for domestic animals to ensure proper levels of iron in their systems; and
- In the manufacture of magnets and magnetic materials.

Exposure to iron(III) oxide dust can cause irritation of the eyes and throat. Long-term exposure to dust particles can cause chronic inflammation of the lungs.

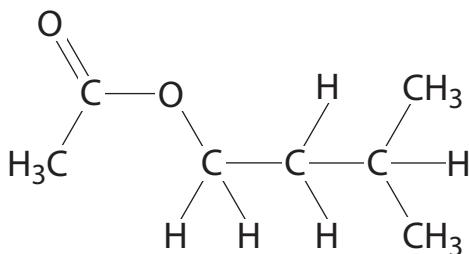
FOR FURTHER INFORMATION

Crutchfield, Charlie. "Re: What Do the Different Colors of Rust Mean, Chemically." MadSci Network.
http://www.madsci.org/posts/archives/2002_03/1015309769.Ch.r.html (accessed on October 12, 2005).

"Ferric Oxide." International Labour Organization.
http://www.ilo.org/public/english/protection/safework/cis/products/icsc/dtasht/_icsc15/icsc1577.htm (accessed on October 12, 2005).

Ricketts, John A. "How a Blast Furnace Works." American Iron and Steel Institute.
<http://www.steel.org/AM/Template.cfm?Section=Home&template=/CM/HTMLDisplay.cfm&ContentID=5433>
(accessed on October 12, 2005).

See Also Carbon Monoxide; Iron(II) Oxide



OTHER NAMES:
Isopentyl acetate;
isoamyl ethanoate;
amylacetic ester

FORMULA:
 $\text{CH}_3\text{COOCH}_2\text{CH}_2\text{CH}(\text{CH}_3)_2$

ELEMENTS:
Carbon, hydrogen,
oxygen

COMPOUND TYPE:
Ester (organic)

STATE:
Liquid

MOLECULAR WEIGHT:
130.18 g/mol

MELTING POINT:
−78.5°C (−109°F)

BOILING POINT:
142.5°C (288.5°F)

SOLUBILITY:
Slightly soluble in
water; soluble in most
organic solvents

KEY FACTS

OVERVIEW

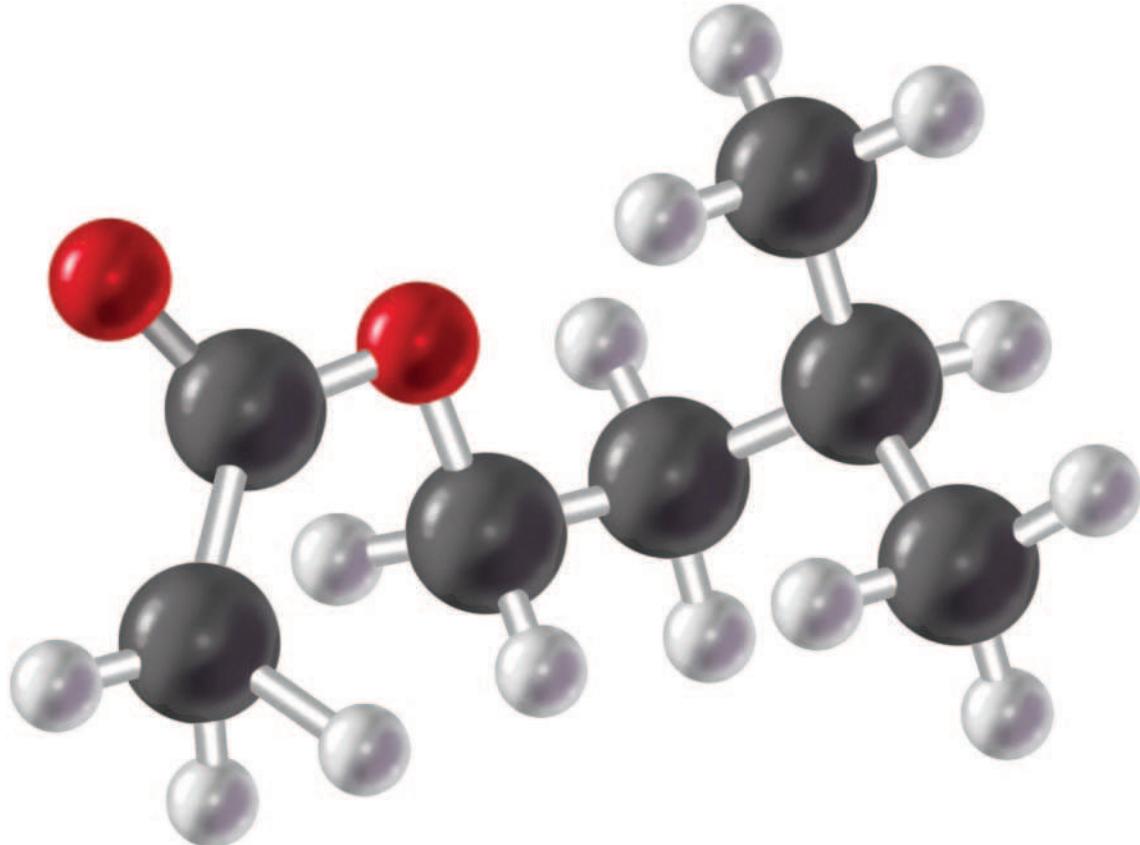
Isoamyl acetate (EYE-so-A-mil AS-uh-tate) is a clear, colorless liquid with a pleasant fruity odor and taste reminiscent of pears or bananas. When prepared for industrial or commercial use, it is often known as pear oil or banana oil.

HOW IT IS MADE

Isoamyl acetate is made commercially by reacting acetic acid (CH_3COOH) with amyl alcohol ($\text{C}_4\text{H}_9\text{CH}_2\text{OH}$) to produce amyl acetate, of which there are eight isomers. The isomers are then separated from each other by fractional distillation to obtain the one desired form, isoamyl acetate.

COMMON USES AND POTENTIAL HAZARDS

Isoamyl acetate is a very popular additive for imparting a pleasant odor or taste to commercial products. Since 1976, the U.S. Patent Office has issued 1,174 patents for inventions



Isoamyl acetate. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. Gray sticks indicate double bonds.

PUBLISHERS RESOURCE GROUP

that contain the compound. Some of the products to which it is added include:

- Foods and drinks, such as gum, candy, mineral water, beer, and syrups used to make soft drinks;
- Personal care products, such as perfume, nail polish, leather polish, and shoe polish;
- Furniture polish, varnishes, and lacquers; and
- Dry cleaning preparation.

In addition to these applications, isoamyl acetate has a number of other uses, including

- The fermentation of grain to produce whiskey;
- To mask unpleasant odors;
- In the manufacture of a number of industrial and household products, including bath sponges, artificial

Interesting Facts

- When a bee stings, it leaves behind traces of isoamyl acetate at the site of the sting. The isoamyl acetate then attracts other bees to the same site, accounting for the tendency for an individual to receive multiple stings at the same point on his or her body.
- One species of Japanese honeybees defends itself from attacks by hornet predators by surrounding the hornet with a ball that consists primarily of isoamyl acetate. The ball becomes so hot that the hornet dies.

leather, artificial silk, rayon, artificial pearls, artificial glass, bronzing fluid, metallic paint, fluorescent lamps, and photographic film;

- In the dyeing and finishing of textiles; and
- As a solvent for old oil paints.

Isoamyl acetate is flammable and is rated as a severe fire hazard. It is also explosive. The chemical should not be used around open flames or sparks. People should not smoke around isoamyl acetate.

Isoamyl acetate is also an irritant to the skin, eyes, and respiratory and digestive systems. If swallowed it may cause sore throat, nausea, and abdominal pain. People who work

Words to Know

DISTILLATION A process of separating two or more substances by boiling the mixture of which they are composed and condensing the vapors produced at different temperatures.

ISOMER One of two or more forms of a chemical compound with the same molecular formula, but different structural formulas and different chemical and physical properties.

with the pure compound are at greater risk for harm from the compound than are those who use it in commercial products.

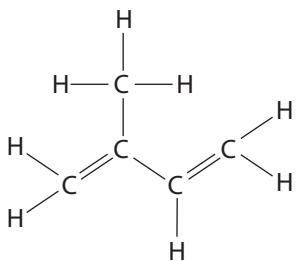
FOR FURTHER INFORMATION

"Isoamyl Acetate." International Programme on Chemical Safety.
<http://www.inchem.org/documents/icsc/icsc/eicso356.htm>
(accessed on October 12, 2005).

"Occupational Safety and Health Guideline for Isoamyl Acetate."
Occupational Safety & Health Administration.
<http://www.osha.gov/SLTC/healthguidelines/isoamylacetate/recognition.html> (accessed on October 12, 2005).

Shukla, Nutan. "Honeybees Come to Know of Queen's Death through Smell." *The Tribune Spectrum*.
<http://www.tribuneindia.com/2003/20030601/spectrum/nature.htm> (accessed on October 12, 2005).

"Unusual Thermal Defence by a Honeybee against Mass Attack by Hornets." *Nature* (September 28, 1995): 334–336.



OTHER NAMES:

2-methyl-
1,3-butadiene

FORMULA:

$\text{CH}_2=\text{CH}(\text{CH}_3)\text{CH}=\text{CH}_2$

ELEMENTS:

Carbon, hydrogen

COMPOUND TYPE:

Alkadiene;
unsaturated
hydrocarbon
(organic)

STATE:
Liquid

MOLECULAR WEIGHT:
68.12 g/mol

MELTING POINT:
−145.9°C (−230.6°F)

BOILING POINT:
34.0°C (93.2°F)

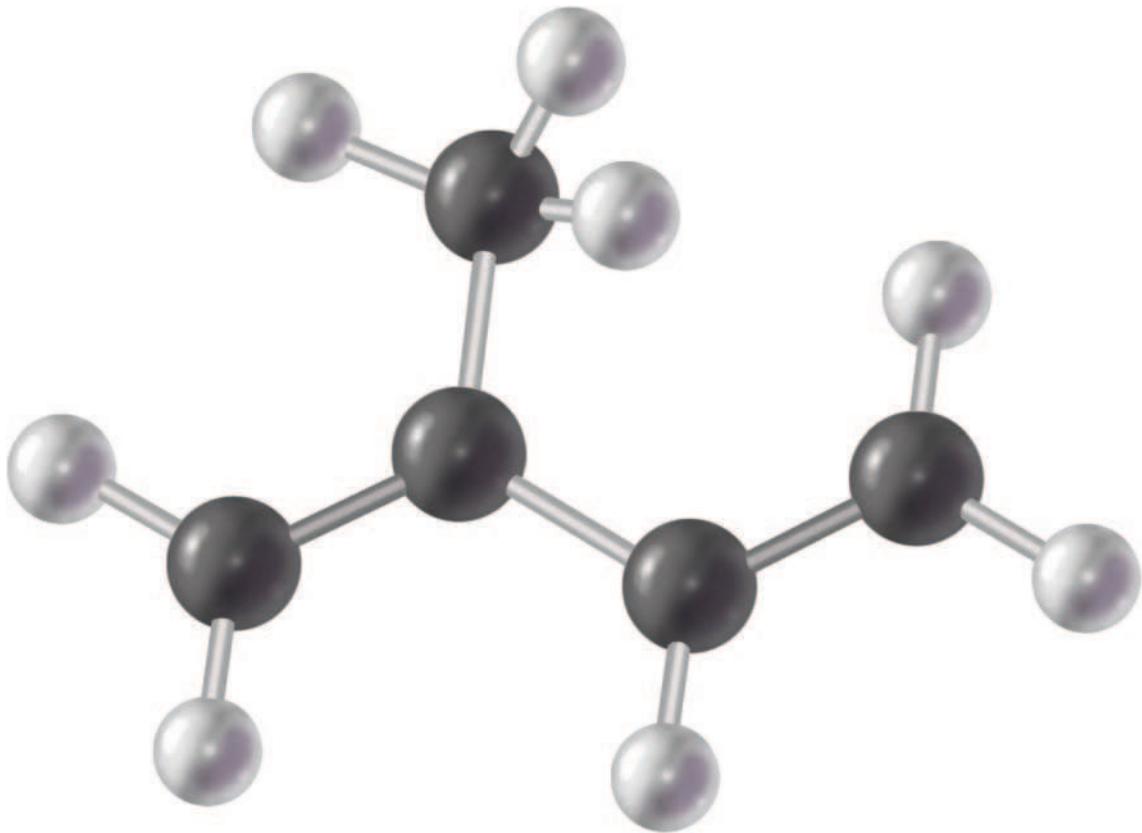
SOLUBILITY:
Insoluble in water;
miscible with ethyl
alcohol, acetone,
ether, and benzene

K E Y F A C T S

OVERVIEW

Isoprene (EYE-so-preen) is a clear, colorless, volatile liquid that is both very flammable and quite explosive. It is classified as a *diene* compound because its molecules contain two (“di-”) double bonds (“-ene”). It is also a member of the terpene family. The terpenes are a large family of organic compounds that contain two or more isoprene units. An example of a terpene is vitamin A, whose molecular formula is $\text{C}_{20}\text{H}_{30}\text{O}$. Vitamin A contains four isoprene units. The terpenes occur abundantly in nature in both plants and animals.

Some common terpenes include geraniol, found in geraniums; limonene, oil of orange; α-pinene, or oil of turpentine; α-farnesene, oil of cintronella; zingiberene, oil of ginger; farnesol, found in lily of the valley; β-selinene, oil of celery; and caryophyllene, oil of cloves. Isoprene is also produced in animal bodies and is said to be the most common hydrocarbon present in the human body. By one estimate, a 70-kilogram (150-pound) person produces about 17 milligrams of



Isoprene. Black atoms are carbon; white atoms are hydrogen. Gray sticks show double bonds; white sticks show single bonds. PUBLISHERS RESOURCE GROUP

isoprene per day. Probably the best-known source of isoprene is natural rubber, which is a polymer consisting of long chains of isoprene units joined to each other.

HOW IT IS MADE

A number of methods are available for preparing isoprene from petroleum. Perhaps the most common process is the cracking of hydrocarbons present in the naphtha portion of refined petroleum. Cracking is the process by which large hydrocarbons are broken down into smaller hydrocarbons either with heat or over a catalyst, or by some combination of heat and catalyst. The naphtha portion of petroleum consists of hydrocarbons with boiling points between about 50°C and 200°C (120°F and 400°F). Other methods for the

Interesting Facts

- Isoprene and other terpenes are now known to undergo reactions that contribute to the development of pollutants, such as ozone and oxides of nitrogen in the atmosphere.
- Isoprene is a key intermediary in the synthesis of cholesterol in the human body.
- The production of isoprene by plants seems to be associated with the process of photosynthesis and is affected by temperature, sunlight, other gases, and other factors.
- The polymer of isoprene is called polyisoprene. It exists in two forms, *cis*- and *trans*-polyisoprene. The two forms are called *geometric isomers*. They have the same kind and number of atoms, but the atoms are arranged differently in the two forms. Natural rubber consists of *trans*-polyisoprene, while another product found in rubber plants, gutta percha, is made of *cis*-polyisoprene.

preparation of isoprene include the dehydrogenation (removal of hydrogen) of isopentene ($\text{CH}_3\text{CH}(\text{CH}_3)\text{CH}=\text{CH}_2$), the pyrolysis (decomposition by high heat) of methylpentene ($\text{CH}_2=\text{C}(\text{CH}_3)\text{CH}_2\text{CH}_2\text{CH}_3$), or the dehydration (removal of water) of methylbutenol ($\text{CH}_3\text{C}(\text{CH}_3)(\text{OH})\text{CH}_2\text{CH}_3$).

COMMON USES AND POTENTIAL HAZARDS

Natural rubber has been known to humans for hundreds of years. Archaeologists have found that the Indians of South and Central America were making rubber products as early as the eleventh century. Until the end of the nineteenth century, natural supplies of rubber obtained from the rubber tree, *Hevea brasiliensis*, were sufficient to meet consumer demand for the product. However, with the development of modern technology—especially the invention of the automobile—natural supplies of the product proved to

be insufficient to meet growing demand. Chemical researchers began to look for ways of producing synthetic forms of rubber.

One approach was to attempt making synthetic rubber with exactly the same chemical composition as that of natural rubber, that is, a polymer of *trans*-polyisoprene. As early as the 1880s, British chemist Sir William Augustus Tilden (1842–1926) was successful in achieving this objective. Tilden found that he could make isoprene by heating turpentine ($C_{10}H_{16}$). The isoprene then polymerized easily when exposed to light. After more than twenty years of research, however, Tilden decided that synthetic *trans*-polyisoprene could never be made economically, and he encouraged his friends to forget about the process.

Over the years, chemists did find ways of making other types of synthetic rubber, and some never abandoned the effort to make synthetic *trans*-polyisoprene. The critical breakthrough needed in this research occurred in about 1953 when Swiss chemist Karl Ziegler (1898–1973) and Italian chemist Giulio Natta (1903–1979) each found a way of polymerizing isoprene in such a way that its geometric structure matched that of natural rubber exactly. A year later, chemists at two of the largest rubber companies in the world, B. F. Goodrich and Firestone, announced that they had developed methods for making synthetic *trans*-polyisoprene using essentially the methods developed earlier by Ziegler and Natta.

In the early twenty-first century, more than 95 percent of the isoprene produced is used to make *trans*-polyisoprene synthetic rubber. The remaining 5 percent is used to make other types of synthetic rubber and other kinds of polymers. A small amount of the compound is used as a chemical intermediary, a substance from which other organic chemicals is made.

Isoprene is a dangerous fire hazard. It also poses a risk to human health and that of other animals. It is an irritant to skin, eyes, and the respiratory system. Upon exposure, it produces symptoms such as redness, watering, and itching of the eyes and itching, reddening, and blistering of the skin. If inhaled, it can irritate the lungs and respiratory system. Isoprene is a known carcinogen.

Words to Know

CARCINOGEN A chemical that causes cancer in humans or other animals.

MISCIBLE Able to be mixed; especially applies to the mixing of one liquid with another.

PHOTOSYNTHESIS The process by which green plants and some other organisms use the energy in sunlight to convert

carbon dioxide and water into carbohydrates and oxygen.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

VOLATILE Able to turn to vapor easily at a relatively low temperature.

FOR FURTHER INFORMATION

“Hazardous Substance Fact Sheet: Isoprene.” New Jersey Department of Health and Senior Services.

<http://www.state.nj.us/health/eoh/rtkweb/1069.pdf> (accessed on December 29, 2005).

“Isoprene.” Shell Chemicals.

<http://www.shellchemicals.com/isoprene/1,1098,1116,00.html> (accessed on December 29, 2005).

“Material Safety Data Sheet: Isoprene MSDS.” ScienceLab.com.

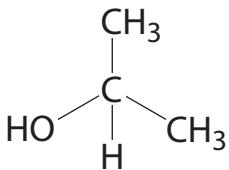
http://www.sciencelab.com/xMSDS_Isoprene_9924409 (accessed on December 29, 2005).

“United States Synthetic Rubber Program, 1939–1945.” National Historic Chemical Landmarks, American Chemical Society.

http://acswebcontent.acs.org/landmarks/landmarks/rbb/rbb_begin.html (accessed on December 29, 2005).

Weissermel, Klaus, and Hans Jürgen Arpe. *Industrial Organic Chemistry*. Weinheim, Germany: Wiley VCH, 2003, 117–122.

See Also Poly(Styrene-Butadiene-Styrene)



OTHER NAMES:
2-propanol;
isopropanol; rubbing
alcohol

FORMULA:
 $\text{CH}_3\text{CHOHCH}_3$

ELEMENTS:
Carbon, hydrogen,
oxygen

COMPOUND TYPE:
Alcohol (organic)

STATE:
Liquid

MOLECULAR WEIGHT:
60.10 g/mol

MELTING POINT:
−87.9°C (−126°F)

BOILING POINT:
82.3°C (180°F)

SOLUBILITY:
Miscible with water
and most common
organic solvents

KEY FACTS

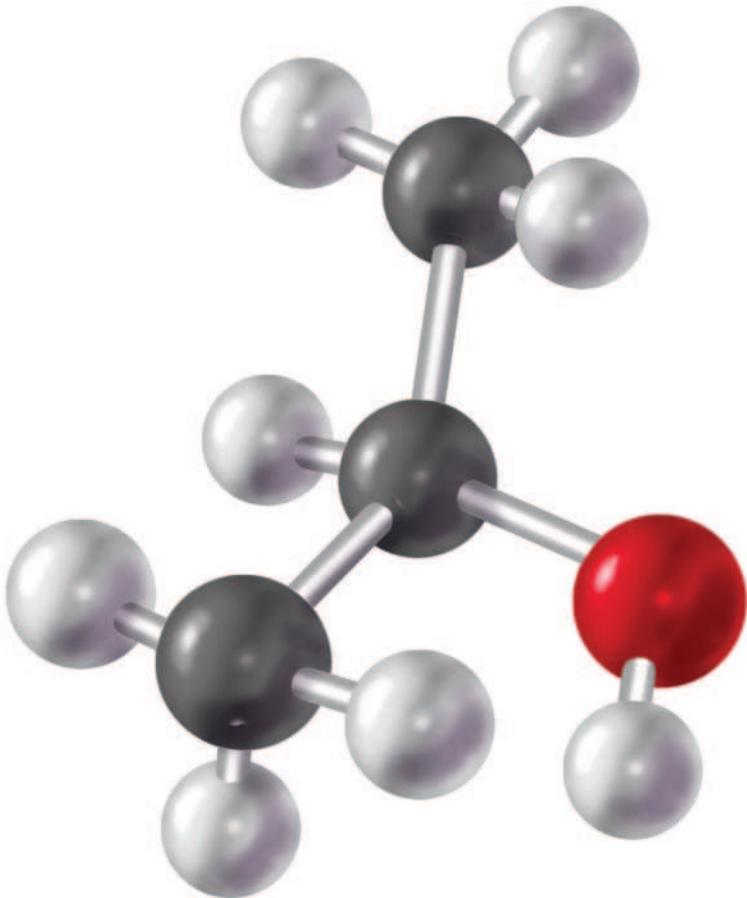
OVERVIEW

Isopropyl alcohol (EYE-so-PRO-pil AL-ko-hol) is a colorless flammable liquid with a sweet odor. In 2004, about 600 million kilograms (about 1.3 billion pounds) of isopropyl alcohol were produced in the United States, with about half of that used as an industrial solvent and about a third used in the preparation of other chemical compounds. It is perhaps best known to many people as rubbing alcohol, usually a 70 percent solution of isopropyl alcohol in water. The compound is commonly used to clean a person's skin before an injection is given. It kills bacteria on the skin and prevents infection.

HOW IT IS MADE

The most popular industrial method for preparing isopropyl alcohol was invented in 1920 by researchers at the Standard Oil Company (now Exxon). In that process, propene (propylene; $\text{CH}_2\text{CH}=\text{CH}_2$) is treated with hydrolyzed with sulfuric acid as a catalyst.

Isopropyl alcohol. Red atom is oxygen; white atoms are hydrogen; and black atoms are carbon. PUBLISHERS RESOURCE GROUP



COMMON USES AND POTENTIAL HAZARDS

Isopropyl alcohol dissolves many other organic compounds easily, so it finds wide use as a solvent for gums; essential and other kinds of oils; alkaloids; certain types of plastics; derivatives of cellulose; paints, varnishes, shellacs, and other types of coatings; and quick-drying inks. Essential oils are oils extracted from plants that have therapeutic value. Alkaloids are organic bases that contain the element nitrogen.

The synthesis of many important organic compounds begins with isopropyl alcohol as a raw material. Among these compounds are acetone, glycerol, and isopropyl acetate, itself widely used as a solvent for organic substances. Among the other uses to which isopropyl alcohol is put are:

Words to Know

HYDROLYSIS The process by which a compound reacts with water to form two new compounds.

MISCIBLE Able to be mixed; especially applies to the mixing of one liquid with another.

- In household and personal care products, such as perfumes, hair dye rinses, nail polishes, shampoos, and after-shave lotions;
- In cleaning products, such as disinfectant soaps and hand and body lotions;
- In antifreezes and as a deicing agent;
- As a coolant in the manufacture of beers; and
- As a preservative for biological specimens.

FOR FURTHER INFORMATION

“Isopropanol.” Spectrum Laboratories.

<http://www.speclab.com/compound/c67630.htm> (accessed on October 12, 2005).

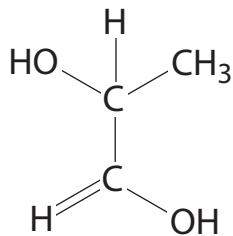
“Isopropyl Alcohol.” New Jersey Department of Health and Senior Services.

<http://www.state.nj.us/health/eoh/rtkweb/1076.pdf> (accessed on October 12, 2005).

“2 Propanol.” International Programme on Chemical Safety.

<http://www.inchem.org/documents/sids/sids/67630.pdf> (accessed on October 12, 2005).

See Also Propylene



KEY FACTS	
OTHER NAMES:	2-hydroxypropanoic acid; α -hydroxypropanoic acid; milk acid
FORMULA:	CH ₃ CHOHCOOH
ELEMENTS:	Carbon, hydrogen, oxygen
COMPOUND TYPE:	Carboxylic acid (organic)
STATE:	Liquid
MOLECULAR WEIGHT:	90.08 g/mol
MELTING POINT:	16.8°C (62.2°F)
BOILING POINT:	Not applicable; decomposes upon heating
SOLUBILITY:	Very soluble in water and ethanol; slightly soluble in ether

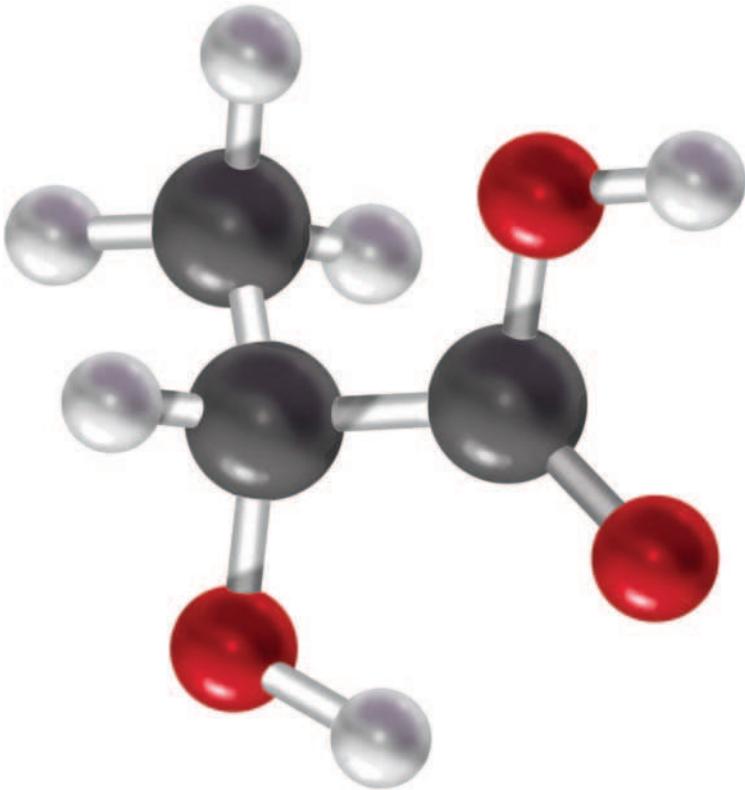
Lactic Acid

OVERVIEW

Lactic acid (LAK-tik AS-id) is a colorless, odorless, syrupy liquid that occurs in two isomeric forms, D-lactic acid and L-lactic acid. Isomers are two or more forms of a chemical compound with the same molecular formula, but different structural formulas and different chemical and physical properties. The D form is produced during metabolic reactions that take place in muscle tissue, while the L form is produced by yeast cells. The synthetic production of lactic acid results in a product consisting of equal amounts of the D and L forms, a mixture known as a racemic mixture.

Lactic acid was first discovered in 1780 by the Swedish chemist Karl Wilhelm Scheele (1742-1786), who called his discovery “acid of milk.” The two isomeric forms of the acid were first identified in 1863 by the German chemist Johannes Wislicenus (1835-1902), and the compound was first produced commercially in 1881 by American chemist Charles E. Avery. Avery patented his invention in 1885 and constructed a factory for the production of lactic acid in Littleton, Massachusetts.

Lactic acid. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP



About 30 million kilograms (72 million pounds) of lactic acid are produced annually in the United States. The most common method of production is the fermentation of glucose by yeast.

HOW IT IS MADE

In muscle cells, lactic acid is the product of anaerobic respiration, the process by which glucose is oxidized in the absence of oxygen to produce energy required by cells. Although some lactic acid is always produced in muscle cells in very low concentrations, it tends to accumulate during exercise, when cells do not receive adequate amounts of oxygen to metabolize oxygen by normal pathways. Lactic acid produced during exercise remains in the body for only short periods of time, sometimes in less than thirty minutes. It is metabolized in the muscle cells where it was produced, resulting in the production of energy, carbon dioxide, water, and other products.

Interesting Facts

- For the better part of a century, athletes and physiologists have considered lactic acid a primary cause of fatigue during high-intensity exercise. However, scientists have learned that lactic acid actually helps to prevent muscle fatigue. Muscle soreness once thought to be caused by lactic acid is, instead, more likely to be a result of damaged muscle cells caused by excess use.
- Lactic acid present on the skin attracts mosquitoes.
- Lactic acid in the body exists in its ionic form, known as lactate.

Lactic acid is also produced by yeast during the process of fermentation. Fermentation is the process by which yeast cells convert glucose to an alcohol and carbon dioxide. Yeast cells use almost precisely the same enzyme in fermentation that muscle cells use in anaerobic respiration. The muscle cell enzyme and the yeast enzyme differ only in the orientation of one group of atoms, resulting in the production of the D isomer in one case and the L isomer in the other.

A synthetic process for the production of lactic acid was first introduced in 1963. That process begins with the addition of hydrogen cyanide (HCN) to acetaldehyde (ethanal; CH_3CHO), resulting in the formation of lactonitrile ($\text{CH}_3\text{CH}_2\text{OCN}$). The lactonitrile is then hydrolyzed, using a strong acid, such as sulfuric acid, as a catalyst, to make lactic acid.

COMMON USES AND POTENTIAL HAZARDS

The primary use for lactic acid in the United States is as a food additive, where it acts as an acidulant and a flavor additive. An acidulant is a compound that provides an acidic environment for foods, as is the case with yogurt, buttermilk, sauerkraut, green olives, pickles, and other acidic foods. As a flavor additive, it adds a tart or tangy flavor to foods and beverages, as well as acting as a preservative to keep them from spoiling. Lactic acid also has a number of important industrial uses, the most important of which is the

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

ELECTROPLATING Adding a layer of nickel, silver, or gold, on another type of metal using an electric current.

FERMENTATION The process by which yeast convert glucose to an alcohol and carbon dioxide.

HYDROLYSIS The process by which a compound reacts with water to form two new compounds.

ISOMER One of two or more forms of a chemical compound with the same molecular formula, but different structural

formulas and different chemical and physical properties.

METABOLISM A process that includes all of the chemical reactions that occur in cells by which fats, carbohydrates, and other compounds are broken down to produce energy and the compounds needed to build new cells and tissues.

MORDANT A substance used in dyeing and printing that reacts chemically with both a dye and the material being dyed to help hold the dye permanently to the material.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

production of other organic chemicals, especially ethyl lactate, acrylic acid, propylene glycol, and the polymer known as polyactide. Polyactide is used in the manufacture of plastic film, fiber, packaging material, and filling materials. Other commercial and industrial applications of lactic acid include:

- As a mordant in dyeing;
- As a solvent for dyes that are not soluble in water;
- For the treatment of animal hides in the preparation of leather products;
- As a catalyst in the production of certain types of plastics; and
- As an additive in electroplating baths.

Lactic acid in normal concentrations poses no safety or health hazards to humans or other animals. One health consequence related to lactic acid, however, is a condition known as gout, a type of arthritis that causes severe pain in the joints. Gout is caused by an accumulation of uric acid in the

blood. Since lactic acid blocks the elimination of uric acid from the body, individuals with excess lactic acid buildup, usually caused by high alcohol consumption, may develop an excess of uric acid crystals in the blood and joints, leading to gout.

FOR FURTHER INFORMATION

“Cell Respiration.” SparkNotes.

<http://www.sparknotes.com/testprep/books/sat2/biology/chapter6section1.rhtml> (accessed on October 14, 2005).

Drake, Geoff. “The Lactate Shuttle Contrary to What You’ve Heard, Lactic Acid Is Your Friend.” *Bicycling* (August 1992): 36.

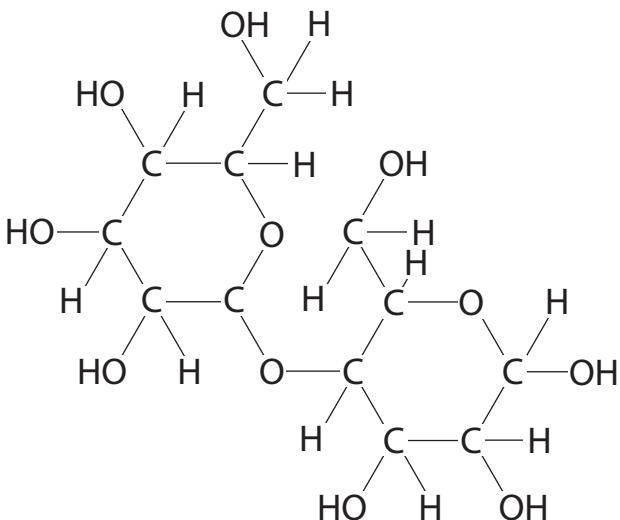
Friel, Joel. “All Athletes: Lactic Acid’s Bad Rap.” Ultrafit’s e Tips For Endurance Athletes. October 2004, Vol. 7, No. 10.

<http://www.ultrafit.com/newsletter/october04.html#Joe> (accessed on October 14, 2005).

“Lactic Acid.” J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/lo522.htm> (accessed on October 14, 2005).

Rogers, Palmer, Jiann Shin Chen, and Mary Jo Zidwick. *Organic Acid and Solvent Production, Part I: Acetic, Lactic, Gluconic, Succinic, and Polyhydroxyalkanoic Acids*. Section 2: Lactic Acid. Available online at http://141.150.157.117:8080/prokPUB/chaphtm/306/04_00.htm (accessed on October 14, 2005).



OTHER NAMES:

D-lactose; milk sugar; many others

FORMULA:

$C_{12}H_{22}O_{11}$

ELEMENTS:

Carbon, hydrogen, oxygen

COMPOUND TYPE:

Disaccharide; carbohydrate (organic)

STATE:
Solid

MOLECULAR WEIGHT:
342.30 g/mol

MELTING POINT:
222.8°C (433.0°F)

BOILING POINT:
Not applicable; decomposes

SOLUBILITY:
Very soluble in water; slightly soluble in ethyl alcohol; insoluble in organic solvents

K E Y F A C T S

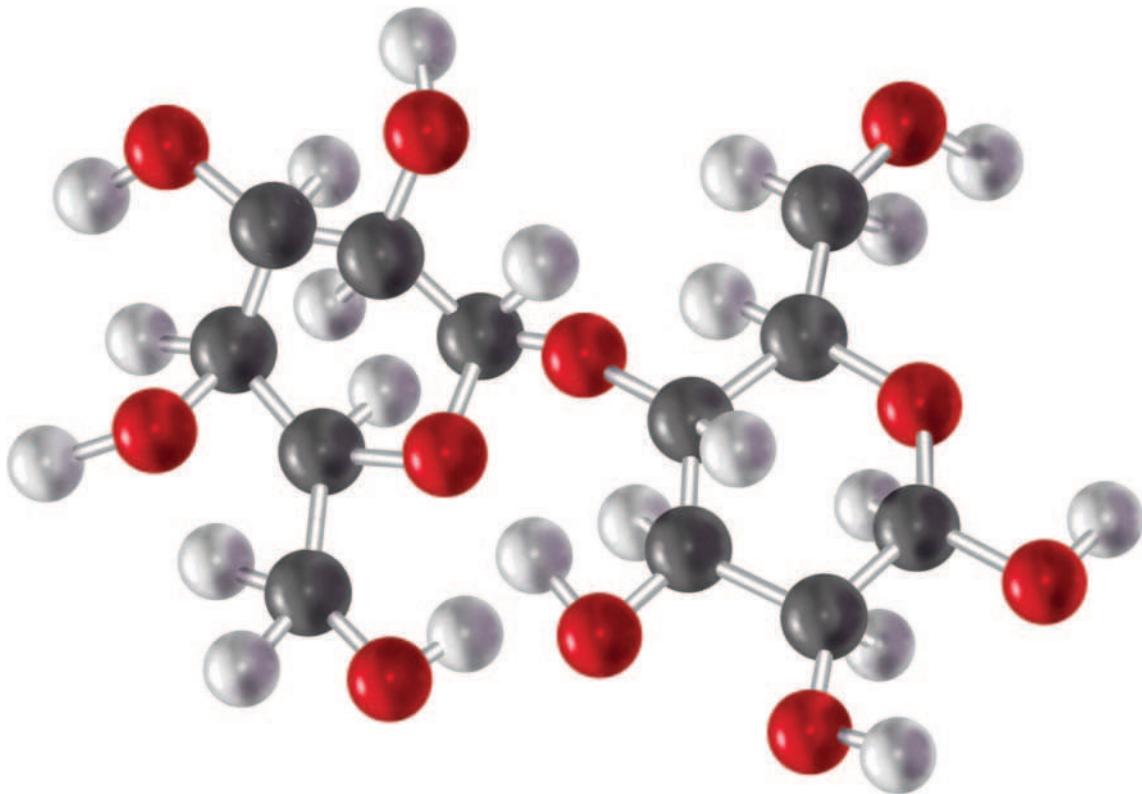
Lactose

OVERVIEW

Lactose (LAK-tose) is a white, odorless, sweet-tasting solid commonly known as milk sugar because it occurs in the milk of many animals, primarily the mammals. The lactose content of milk ranges from about 2 to 8 percent in cows and 5 to 8 percent in human milk. Lactose occurs in two isomeric forms, α -lactose and β -lactose, with the latter somewhat sweeter than the former. The alpha form tends to occur as the monohydrate, $C_{12}H_{22}O_{11} \cdot H_2O$.

HOW IT IS MADE

Lactose is synthesized in the mammary (milk-producing) glands of mammals. The milk of such animals contains an enzyme called lactose synthetase, which acts on the compound uridine diphosphate D-galactose to produce lactose. The compound is obtained commercially from whey, a byproduct of the cheese-making process. The solids in whey contain about 70 percent lactose by weight. These solids are



Lactose. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. PUBLISHERS RESOURCE GROUP

filtered to remove the proteins they contain. After removal of minerals in the whey, the resulting solution consists of about 50 to 65 percent by weight, which is allowed to crystallize out of the resulting solutions. The lactose produced by this process is a racemic mixture of the D and L isomers. To obtain the alpha isomer, the lactose is dissolved with water and treated with activated carbon to remove any color. When the water evaporates from the solution, α -lactose monohydrate remains. The product is most commonly made available in this form. To obtain the beta isomer, α -lactose is heated with water in the presence of a base, which converts the alpha isomer to its beta form.

COMMON USES AND POTENTIAL HAZARDS

Both forms of lactose are used in the preparation of baby foods and food for infants and for convalescents. It is also used in the dairy industry to feed foals that have been

Interesting Facts

- Only one group of mammals does not produce lactose in their milk, the Pinnipedia, a group that includes seals, sea lions, and walruses.
- Alpha lactose is found in cow's milk and beta lactose in human milk. During pasteurization of cow's milk, the alpha isomer is converted into the beta isomer.

orphaned. Lactose is also used as a food additive, primarily as a humectant and to increase the sweet flavor of a food. A humectant helps foods retain their moisture. It is used in products such as ice creams, baked goods, confectionary items, whipped toppings, and breakfast foods. Lactose may also be added to a number of foods with limited natural sweetness, such as margarine, butter, frozen vegetables, and processed meats. It is also used as a filler for many pharmaceutical products. A filler acts bulk to the product without significantly changing its nutritional or medical properties.

Many people are lactose intolerant. Lactose intolerance is a condition that develops when a person lacks enough (or any) of the enzyme lactase that is responsible for digestion of lactose in the body. People who are lactose intolerant and consume lactose experience a range of unpleasant symptoms, including fluid retention, gas, cramps, and diarrhea. By some estimates, up to 75 percent of the world's population may experience lactose intolerance to a greater or lesser degree.

One method for dealing with lactose intolerance is for a person to avoid consumption of any food or food product that contains lactose. The option has been made somewhat easier by the introduction of lactose-free products by a number of food companies. Another method to avoid the problems associated with lactose intolerance is to use a dietary supplement that contains lactase, restoring to the body the enzyme that it lacks naturally.

Words to Know

HYDRATE A chemical compound formed when one or more molecules of water is added physically to the molecule of some other substance.

MAMMALS A group of warm-blood animals whose females produce milk for their

young. This group includes humans, cows, bears, and dogs, among others.

RACEMIC MIXTURE A mixture of equal amounts of two opposite isomers (the “right-handed,” or “D” form and the “left-handed,” or “L” form).

FOR FURTHER INFORMATION

“Lactose.” J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/l1044.htm>
(accessed on October 14, 2005).

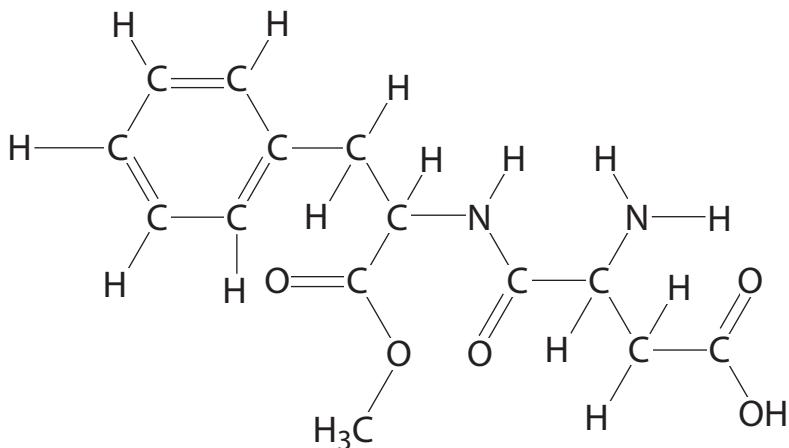
“Lactose Intolerance.” Medline Plus.

<http://www.nlm.nih.gov/medlineplus/lactoseintolerance.html>
(accessed on October 14, 2005).

“Lactose Intolerance.” National Digestive Diseases Information Clearinghouse.

<http://digestive.niddk.nih.gov/ddiseases/pubs/lactoseintolerance/> (accessed on October 14, 2005).

See Also Sucrose



OTHER NAMES:

Aspartame

FORMULA:

C₁₄H₁₈N₂O₅

ELEMENTS:

Carbon, hydrogen, nitrogen, oxygen

COMPOUND TYPE:

Ester (organic)

STATE:

Solid

MOLECULAR WEIGHT:

294.30 g/mol

MELTING POINT:

246.5°C (475.7°F)

BOILING POINT:

Not applicable

SOLUBILITY:

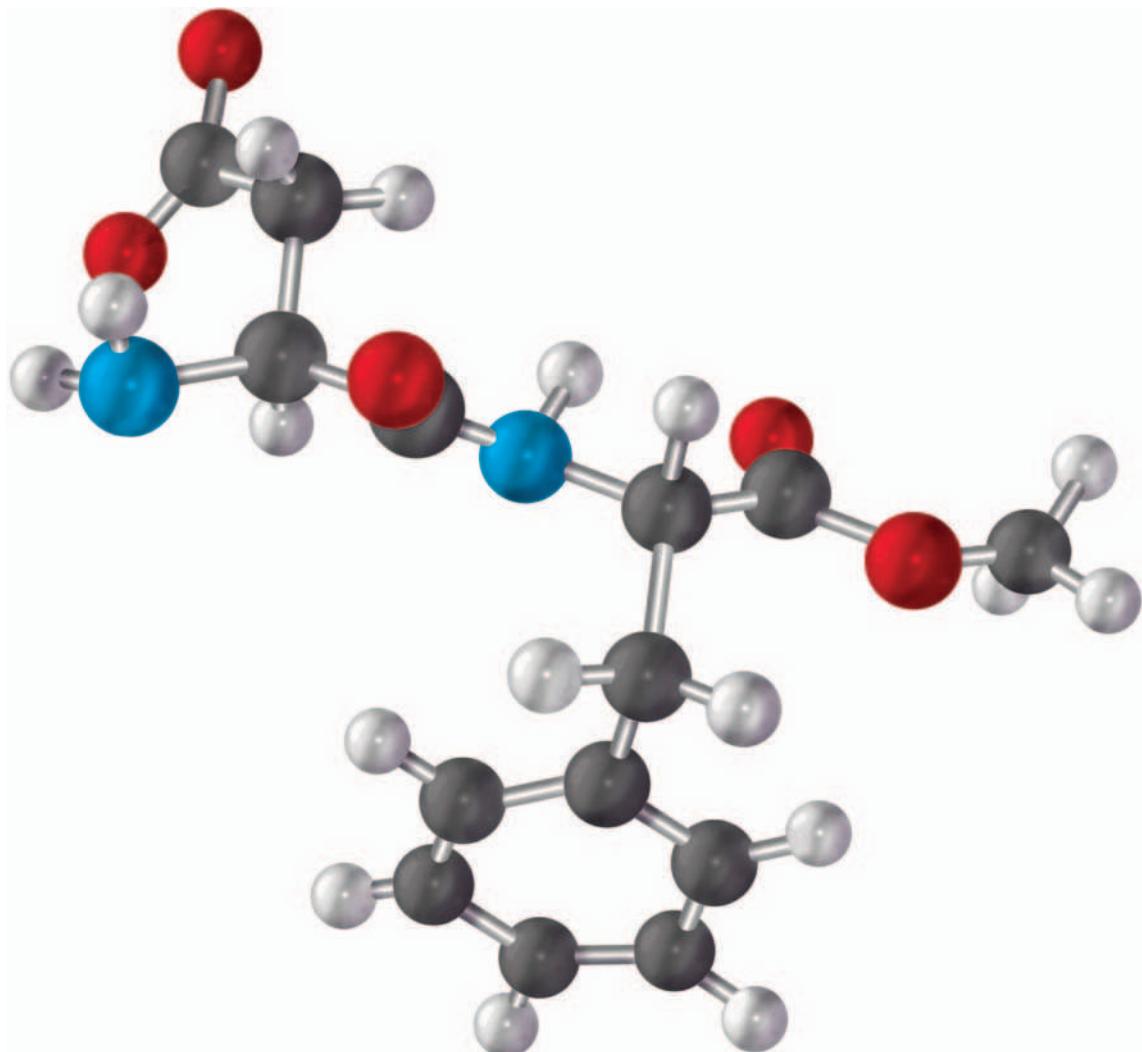
Slightly soluble in water; insoluble in alcohol, benzene, ether, and most other organic solvents

KEY FACTS

OVERVIEW

L-aspartyl-L-phenylalanine methyl ester (ell-ass-par-TEEL ell-fee-no-AL-uh-neen METH-el ESS-ter) is an artificial sweetener more widely known as aspartame. It is sold under a number of brand names, including NutraSweet®, Equal®, Spoonful®, Benevia®, Indulge®, NatraTaste®, and Equal-Measure®. Unlike sugar, which is a carbohydrate, aspartame is a dipeptide, a compound made of two amino acids joined to each other. It is 180 to 200 times as sweet as sucrose (table sugar), but provides no calories to a person's diet. It is a satisfactory substitute for sugar, therefore, for people who must or wish to reduce their caloric intake.

Aspartame was discovered accidentally in 1965 by James M. Schlatter, an employee of the G. D. Searle pharmaceutical company. Schlatter was searching for a chemical compound that could be used to treat ulcers. L-aspartyl-L-phenylalanine methyl ester was one of the compounds he made during his investigations. He accidentally got some of the compound on his fingers, a fact that he did not notice until later in the day.



L Aspartyl L phenylalanine methyl ester. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atoms are nitrogen. Gray sticks show double bonds. PUBLISHERS RESOURCE GROUP

Then, when he licked his fingers to pick up a piece of paper, he noticed the intensely sweet taste of the compound. To confirm his suspicions about the compound, he added some of it to his coffee—a practice that violates all laboratory safety rules!—and found that it was more than satisfactory as a substitute for sugar. Searle began the testing and application procedure needed to gain approval from the U.S. Food and Drug Administration, a long and drawn-out procedure that ended only in 1981, 25 years after its discovery.

Interesting Facts

- The Aspartame Information Center claims that the compound is now used in more than 6,000 commercial products and consumed by more than 100 million people worldwide.

HOW IT IS MADE

Aspartame is made by a relatively simple procedure in which two amino acids, aspartic acid and phenylalanine, are reacted with each other to form a two-amino-acid product, called a dipeptide. The carboxylic acid group in the dipeptide is then reacted with methanol (methyl alcohol; CH_3OH) to obtain the methyl ester of the compound. One problem that makes the preparation somewhat more difficult is that both aspartic acid and phenylalanine have stereoisomers. The term stereoisomer refers to two forms of a compound that contain the same kind and number of atoms, but differ in the orientation in space ("stereo-") of some of the atoms. Because of these stereoisomers, four different kinds of aspartame are formed during the preparation described above – D & D; L & L; D & L; L & D (the latter two are different from each other). Only one is the desired product, the one that contains only "L" stereoisomers.

COMMON USES AND POTENTIAL HAZARDS

Aspartame's primary commercial use is as an artificial sweetener. Some of the products that contain aspartame are breath mints, carbonated soft drinks, cereals, chewing gum, sugar-free gelatin, hard candies, ice cream toppings, sugar-free ice cream, ready-to-drink iced tea, instant cocoa mix, jams and jellies, juice drinks, nutritional bars, protein nutritional drinks, diet puddings, sugar-free chocolate syrup, sugar-free cookies, table top sweeteners, and fat- and sugar-free yogurt. The compound is also used to a lesser extent as a flavor enhancer, a substance that intensifies an already-existing flavor.

L ASPARTYL L PHENYLALANINE METHYL ESTER

The use of aspartame as a food additive in the United States has a long and complex history. As required for FDA approval, Searle conducted a number of studies to show that aspartame is safe for human consumption. Based on the results of those studies, the FDA first granted approval for the use of aspartame in certain types of foods on July 26, 1974. Less than a month later, two concerned citizens, James Turner and Dr. John Olney, filed a petition objecting to the FDA's decision, citing possible errors in Searle's testing procedures on aspartame. After continued studies extending over a seven-year period, the FDA once again approved aspartame for use in dry foods. A year later, the agency extended its approval for aspartame to dry beverage mixes and add-in sweeteners and, in 1983, to beverages and other foods.

The campaign against aspartame has not lessened over the years. Today, a number of individuals and groups maintain organizations and websites whose purpose it is to have aspartame banned as a food additive. Continued studies by the manufacturer and independent researchers seem to confirm the safety of aspartame, but have found that two groups of people do face some level of risk by ingesting the compound. The first group consists of people who seem to have some natural allergy to the compound and experience symptoms such as anxiety attacks, breathing difficulties, dizziness, unusual fatigue, headaches, heart palpitations, skin rashes, muscle spasms, nausea, numbness, respiratory allergies, weight gain, and memory loss. Individuals who experience such problems are encouraged to avoid eating or drinking products that contain aspartame.

A second group for whom aspartame is a far more serious problem consists of individuals with phenylketonuria. Phenylketonuria is a genetic disorder in which a person's body is unable to metabolize the amino acid phenylalanine, a component of aspartame. If phenylalanine is ingested by someone with phenylketonuria, the amino acid builds up in the body and causes a number of organic problems, including mental retardation, damage to the organs, and muscular disorders. Any product that contains aspartame is now required by law to have a warning label aimed at phenylketonurics so that they will not accidentally ingest the compound.

Words to Know

DIPEPTIDE A compound made of two amino acids joined to each other.

METHYL ESTER A compound formed when methyl alcohol (methanol) reacts with an organic acid.

PHENYLCETONURIA A genetic disorder in which a person's body is unable to metabolize the amino acid phenylalanine.

STEREOISOMER One of a pair of molecules, both of which have the same kinds and number of atoms, but which differ in how some of the atoms are oriented in relation to each other.

FOR FURTHER INFORMATION

"Aspartame." Chemical Land 21.

<http://www.chemicalland21.com/lifescience/foco/ASPARTAME.htm> (accessed on September 21, 2005).

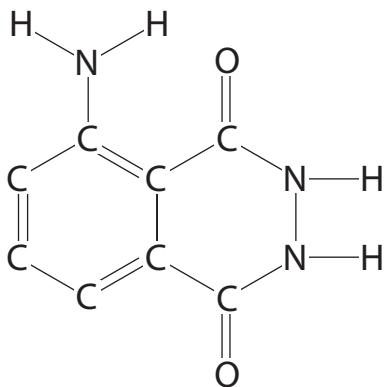
"Aspartame Information Center."

<http://www.aspartame.org/> (accessed on September 21, 2005).

"Low Calorie Sweeteners: Aspartame." Calorie Control Council.

<http://www.caloriecontrol.org/aspartame.html> (accessed on September 21, 2005).

Metcalfe, Ed., et al. "Sweet Talking." *The Ecologist* (June 2000): 16.



OTHER NAMES:

5-amino-2,3-dihydro-1,4-phthalazinedione;
3-aminophthalhydrazide

FORMULA:

C₈H₇N₃O₂

ELEMENTS:

Carbon, hydrogen,
nitrogen, oxygen

COMPOUND TYPE:

Heterocyclic ring
(organic)

STATE:

Solid

MOLECULAR WEIGHT:

177.16 g/mol

MELTING POINT:

319°C to 320°C (606°F
to 608°F)

BOILING POINT:

Not applicable;
decomposes

SOLUBILITY:

Slightly soluble in
water; soluble in
alcohol

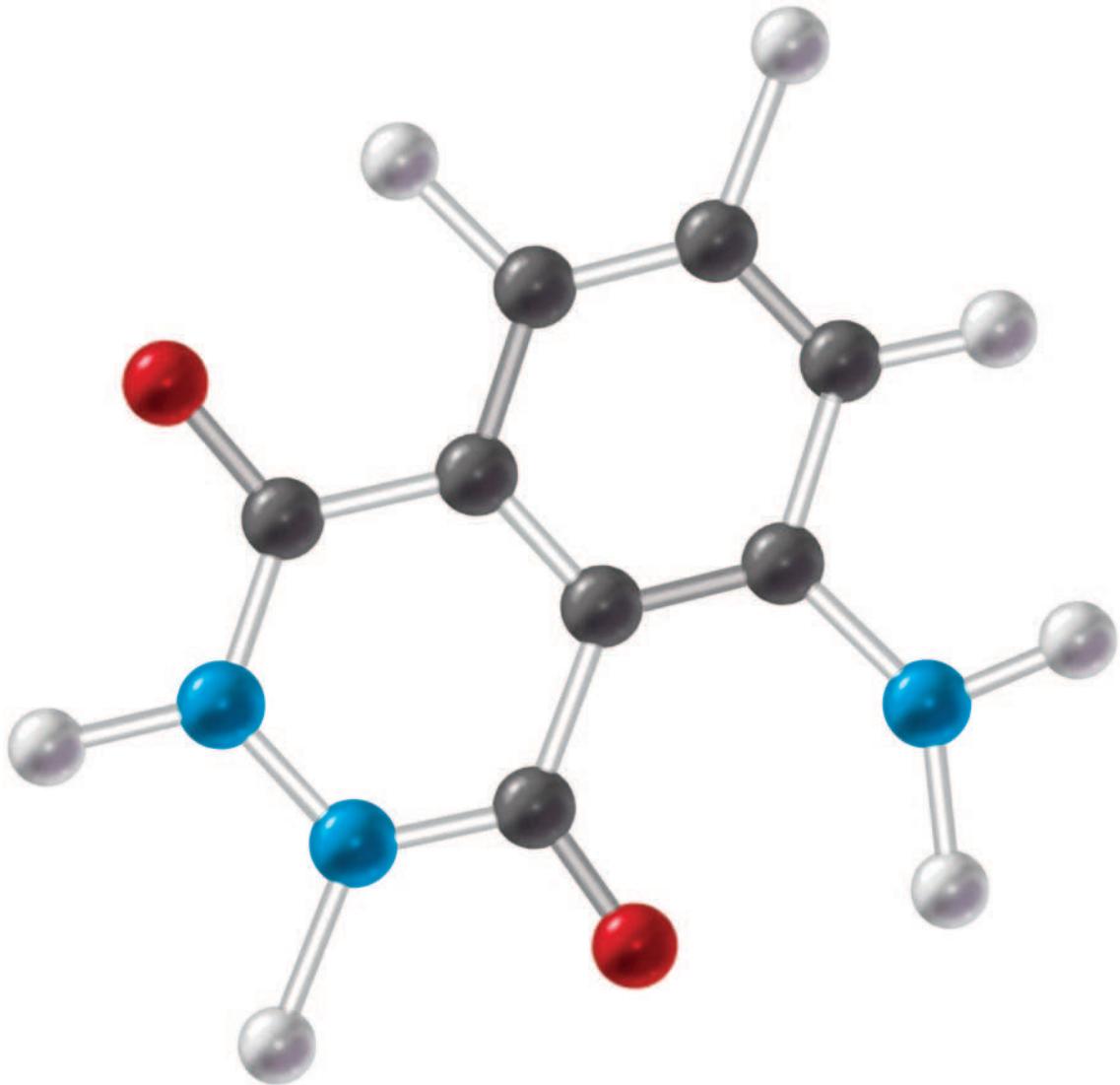
K
E
Y
F
A
C
T
S

Luminol

OVERVIEW

Luminol (LOO-min-ol) is a substance that glows when it comes in contact with blood. It was discovered in the late nineteenth century, but chemists found little use for the compound for half a century. Then, in 1928, the German chemist H. O. Albrecht found that the addition of hydrogen peroxide to luminol produces a bluish-green glow, an example of the phenomenon known as chemiluminescence. Chemiluminescence is the process by which light is emitted as the result of a chemical reaction. Albrecht found that the reaction between hydrogen peroxide and luminol required a catalyst, a small amount of a metal such as copper or iron.

The most important application for luminol was discovered in 1937 by the German forensic scientist Walter Specht (1907–1977), at the University Institute for Legal Medicine and Scientific Criminalistics in Jena, Germany. Specht found that blood itself could act as the catalyst needed to produce chemiluminescence with luminol. He simply sprayed a mixture of luminol and hydrogen peroxide on a



Luminol. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atoms are nitrogen. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

drop of blood, and the blood emitted a bluish-green glow. Specht later determined the explanation for this reaction. Blood contains a protein called hemoglobin that carries oxygen from the lungs to cells. Hemoglobin is a complex molecule with a single iron atom at its center. The small amount of iron in hemoglobin is sufficient to initiate the chemiluminescent reaction between luminol and hydrogen peroxide.

Interesting Facts

- Luminol can be used to detect bloodstains that are many years old.
- One disadvantage of using luminol in testing for blood is that it destroys the sample being investigated, making further tests on the same sample impossible.

HOW IT IS MADE

Luminol is prepared commercially by treating 3-nitrophthalic acid ($C_8H_5NO_6$) with hydrazine (NH_2NH_2), resulting in the formation of nitrophthalhydrazide. Nitrophthalhydrazide is then treated with sodium bisulfite ($NaHSO_3$) to obtain luminol.

COMMON USES AND POTENTIAL HAZARDS

The most common application of luminol is to find traces of blood at crime scenes. When a violent crime is committed, a certain amount of blood is often spilled on the floor, walls, furniture, and other objects at the crime scene. The perpetrator of the crime may attempt to clean up after the crime, but it is virtually impossible to remove all traces of blood. Forensic scientists who investigate a crime scene often assume that blood is present, even if it is not obvious. They check for the blood by spraying a mixture of luminol and hydrogen peroxide around the crime scene. If blood is present, it glows with a bluish-green color. The distribution of the blood can provide information as to where the crime was committed, whether the injured or murdered person was moved, and, if so, in what direction. Investigators typically take photographs of the illuminated crime scene for study at a later date.

False positive results are possible with a luminol test. A false positive test is a test in which the results seem to indicate the presence of blood even if it is not actually there. Some metals, plants, paints, cleaning materials, and other

Words to Know

CATALYST A material that increases the rate of a chemical reaction without

undergoing any change in its own chemical structure.

substances may act as catalysts for the reaction between luminol and hydrogen peroxide and give a false positive test. For this reason, positive tests obtained by the luminol reagent are always subjected to further tests to confirm the results.

Luminol does have applications beyond criminal investigations. It is the active ingredient in glow sticks, the plastic sticks that glow green when broken. Chemists use the compound in chromatography, a process by which the components of a mixture are separated from each other, as well as in studies of DNA patterns and other biochemical tests.

FOR FURTHER INFORMATION

Genge, Ngaire E. *The Forensic Casebook: The Science of Crime Scene Investigation*. New York: Ballantine, 2002.

“Material Safety Data Sheet for Luminol = 3 Aminophthalhydrazide, 98%.” Department of Chemistry, Iowa State University. <http://avogadro.chem.iastate.edu/MSDS/luminol.htm> (accessed on October 14, 2005).

“Nitric Oxide/NOS Detection (incl. Kits/Sets).” Axxora. [http://www.axxora.com/nitric_oxide_nos_detection_\(incl._kits_sets\)_ALX_610_002/opfa.1.1.ALX_610_002.169.4.1.html](http://www.axxora.com/nitric_oxide_nos_detection_(incl._kits_sets)_ALX_610_002/opfa.1.1.ALX_610_002.169.4.1.html) (accessed on October 14, 2005).

“Technical Note: Hemaglow™.” Lightning Powder Company. <http://www.redwop.com/technotes.asp?ID=118> (accessed on October 14, 2005).



OTHER NAMES:	Magnesium dichloride
FORMULA:	MgCl_2
ELEMENTS:	Magnesium, chlorine
COMPOUND TYPE:	Inorganic salt; binary salt
STATE:	Solid
MOLECULAR WEIGHT:	95.21 g/mol
MELTING POINT:	714°C (1,320°F)
BOILING POINT:	1,412°C (2,574°F)
SOLUBILITY:	Soluble in water; moderately soluble in ethyl alcohol

KEY FACTS

Magnesium Chloride

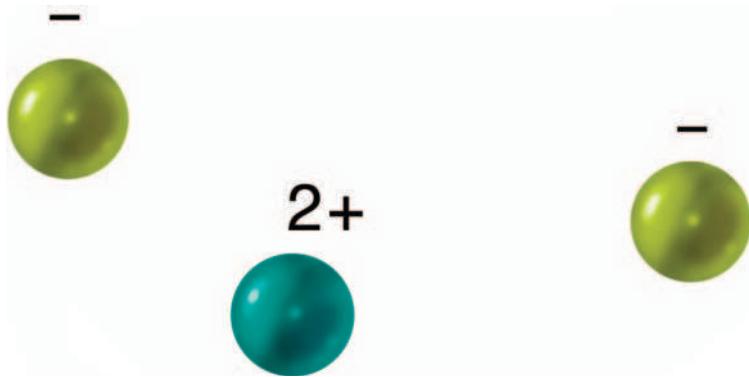
OVERVIEW

Magnesium chloride (mag-NEE-zee-um KLOR-ide) is a white crystalline solid that is strongly deliquescent. It absorbs moisture from the air to become the hydrated form, magnesium chloride hexahydrate ($\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$). A deliquescent substance is one that takes on moisture from the air, often to the extent of dissolving in its own water of hydration. Magnesium chloride is an important industrial chemical, used in the production of magnesium, textile and paper manufacture, and cements; in refrigeration and fireproofing; and as a deicing agent.

HOW IT IS MADE

Magnesium chloride is extracted from seawater or brine, of which it is a component, and from minerals, such as carnallite ($\text{KCl} \cdot \text{MgCl}_2 \cdot \text{H}_2\text{O}$) and bischofite ($\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$). The usual procedure is to treat seawater, brine, or the mineral with lime (CaO), calcined dolomite ($\text{CaO} \cdot \text{MgO}$), or caustic soda (sodium

Magnesium chloride hexahydrate. Green atoms are chlorine; and turquoise atom is magnesium. PUBLISHERS RESOURCE GROUP



hydroxide; NaOH) to make magnesium hydroxide [Mg(OH)₂], which is then treated with hydrochloric acid (HCl) to recover magnesium chloride, which is usually obtained as the crystalline hexahydrate. The pure compound is also produced by heating the double salt, magnesium ammonium chloride (MgCl₂·NH₄Cl·6H₂O), which first loses its water of hydration to form the anhydrous double salt (MgCl₂·NH₄Cl). With further heating, the ammonium chloride sublimes, leaving behind pure anhydrous magnesium chloride.

COMMON USES AND POTENTIAL HAZARDS

The largest single use of magnesium chloride is in the production of magnesium metal. The metal is obtained through the electrolysis of molten magnesium chloride in a process developed by the American chemist and inventor Herbert Henry Dow (1866–1930) in 1916. The Dow process is still the primary method used for the production of magnesium metal. Other commercial and industrial uses of magnesium chloride include:

- In the manufacture of disinfectants;
- In the fireproofing of steel beams, wooden panels, and other materials;
- As a component of fire extinguishers;
- In the manufacture of so-called Sorel cement, a mixture of magnesium chloride and magnesium oxide, also known as oxychloride cement;
- As a binder to control dust on dirt roads;
- As a deicing compound;

Interesting Facts

- Tofu is traditionally prepared by treating soy milk with magnesium chloride or calcium sulfate.

- To remove suspended particles in water and sewage treatment plants;
- For the treatment of cotton and wool fabrics;
- In the processing of sugar beets;
- To keep drilling tools cool; and
- In the manufacture of paper and ceramic materials.

Magnesium chloride is a skin, nose, and eye irritant, although that hazard is usually a matter of concern only to those who work with the pure compound. It is also toxic by ingestion. Swallowing the compound can produce nausea, vomiting, and diarrhea. Inhalation of magnesium chloride fumes can irritate the lungs and respiratory tract, producing a condition known as metal fume fever that resembles the flu.

Words to Know

ANHYDROUS A compound that lacks any water of hydration.

DELIQUESCENT Describing a substance that takes on moisture from the air, often to the extent of dissolving in its own water of hydration.

HYDRATE A chemical compound formed when one or more molecules of water is

added physically to the molecule of some other substance.

SUBLIMATION The process by which a solid changes directly into a gas without first melting.

WATER OF HYDRATION Water that has combined with a compound by some physical means.

FOR FURTHER INFORMATION

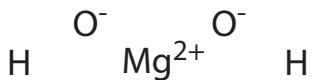
“Magnesium Chloride.” ChemicalLand21.com.

<http://www.chemicaland21.com/arokorhi/industrialchem/inorganic/MAGNESIUM%20CHLORIDE.htm> (accessed on October 14, 2005).

“Magnesium Chloride.” J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/mo156.htm> (accessed on October 14, 2005).

See Also Magnesium Hydroxide

**OTHER NAMES:**

Magnesium hydrate;
milk of magnesia;
magnesia magma

FORMULA:

Mg(OH)_2

ELEMENTS:

Magnesium,
hydrogen, oxygen

COMPOUND TYPE:

Inorganic base

STATE:

Solid

MOLECULAR WEIGHT:

58.33 g/mol

MELTING POINT:

Decomposes at
 350°C (660°F)

BOILING POINT:

Not applicable

SOLUBILITY:

Virtually insoluble in
water and alcohol,
soluble in dilute acids
and solutions of
ammonium salts

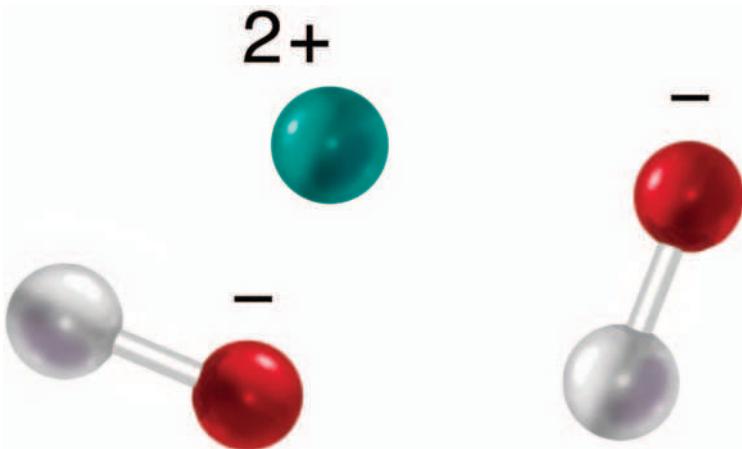
KEY FACTS**OVERVIEW**

Magnesium hydroxide (mag-NEE-zee-um hye-DROK-side) is a white powder with no odor, found in nature as the mineral brucite. Perhaps the best known form of the compound is a milky liquid known as milk of magnesia, a product used to treat upset stomach and constipation. Milk of magnesia was invented in 1817 by the Irish pharmacist Sir James Murray (1788-1871). Murray built a plant to produce a mixture of magnesium hydroxide in water that he sold for the treatment of a variety of ailments, including heartburn, stomach acidity, bladder and bowel problems, and “female problems.” He said that the liquid mixture was much more effective than powdery magnesium hydroxide which had previously been used for the same purposes.

In 1880, New York chemist Charles Henry Phillips (1820-1882) invented the name “milk of magnesia” and opened his own factory for producing the product. The name Phillips Milk of Magnesia is one of the oldest and best known over-the-counter medicines ever made in the United States.

Magnesium hydroxide. Red atoms are oxygen; white atoms are hydrogen; and turquoise atom is magnesium.

PUBLISHERS RESOURCE GROUP



HOW IT IS MADE

Magnesium hydroxide is prepared by reacting a magnesium salt, such as magnesium chloride (MgCl_2), with sodium hydroxide (NaOH). In a similar procedure, seawater (which contains small amounts of magnesium chloride) is treated with lime (calcium oxide; CaO). The water, lime, and magnesium chloride react to produce magnesium hydroxide, which settles out of solution as a precipitate.

COMMON USES AND POTENTIAL HAZARDS

The best known use for magnesium hydroxide is as an antacid in the form of milk of magnesia. Since magnesium hydroxide is a base, it reacts with excess stomach acid, which reduces heartburn and the discomfort of upset stomach.

Interesting Facts

- The element magnesium is named after a region in Greece called Magnesia in ancient times. Today the same region is called Manisa. Compounds of

magnesium, including magnesium hydroxide, were abundant in the region and were given the name of *magnesia lithos*, or “stones of Magnesia.”

Words to Know

ANTACID A medicine used for the treatment of upset stomach, acid indigestion, and related symptoms.

CLARIFIER A substance that removes impurities from another substance.

PRECIPITATE A solid material that settles out of a solution, often as the result of a chemical reaction.

SUSPENSION A mixture of two substances that do not dissolve in each other.

Milk of magnesia also acts as a laxative because it increases the flow of water into the intestines, stimulating a bowel movement. Milk of magnesia consists of an 8 percent suspension of solid magnesium hydroxide in water. Other chemicals, such as calcium carbonate and aluminum hydroxide, are sometimes added to the mixture to increase its effectiveness.

Magnesium hydroxide also has a number of important industrial uses, such as:

- A clarifier (a substance that removes impurities) in the refining of sugar;
- An additive in the treatment of wastewater to neutralize acids present in the wastes;
- A flame retardant coating on fabrics and other materials used by consumers and industries;
- An additive to fuel oils;
- An additive in toothpastes; and
- A drying agent in some food products.

FOR FURTHER INFORMATION

Dean, Carolyn. *The Miracle of Magnesium*. New York: Ballantine Books, 2003.

"Magnesium Hydroxide." Chemical Land 21.

<http://www.chemicaland21.com/arokorhi/industrialchem/inorganic/MAGNESIUM%20HYDROXIDE.htm> (accessed on September 14, 2005).

See Also Magnesium Oxide

OTHER NAMES:	Magnesia; calcined magnesia; magnesia ulba
FORMULA:	MgO
ELEMENTS:	Magnesium, oxygen
COMPOUND TYPE:	Metallic oxide (inorganic)
STATE:	Solid
MOLECULAR WEIGHT:	40.30 g/mol
MELTING POINT:	2,825°C (5,117°F)
BOILING POINT:	3,600°C (6,500°F)
SOLUBILITY:	Slightly soluble in water; soluble in most acids; insoluble in alcohol

KEY FACTS

Magnesium Oxide

OVERVIEW

Magnesium oxide (mag-NEE-see-um OK-side) is available commercially in several forms, depending on the way it is prepared and the use for which it is intended. Most forms can be classified as either “light” or “heavy” depending on particle size, purity, and method of production. It occurs in nature in the form of the mineral periclase. In its purest form, magnesium oxide is a colorless or white crystalline material or very fine powder, with no odor and a bitter taste.

HOW IT IS MADE

A number of methods are available for the preparation of magnesium oxide. Most methods begin with either magnesium hydroxide [Mg(OH)₂] or magnesium carbonate (MgCO₃), either of which is heated under controlled conditions to produce the desired product. The primary methods of preparation are:

- Dead-burning, in which the hydroxide or carbonate is heated to temperatures ranging from 1,500°C to



Magnesium oxide. Black atom is magnesium; red atom is oxygen; the stick represents a double bond. PUBLISHERS RESOURCE GROUP

$2,000^{\circ}\text{C}$ ($3,000^{\circ}$ to $4,000^{\circ}\text{F}$), so-called because the final product is largely chemically unreactive;

- Hard-burning (also known as caustic-burned), in which the hydroxide or carbonate is heated to somewhat lower temperatures, between $1,000^{\circ}\text{C}$ and $1,500^{\circ}\text{C}$ ($2,000^{\circ}\text{F}$ and $3,000^{\circ}\text{F}$), allowing the compound to retain some of its chemical reactivity;
- Light-burning, in which the temperature is kept between 700°C and $1,000^{\circ}\text{C}$ ($1,500^{\circ}\text{F}$ and $2,000^{\circ}\text{F}$), resulting in a product with even more reactivity; and
- Fusion, in which the hydroxide or carbonate is heated to temperatures in excess of $2,650^{\circ}\text{C}$ ($4,800^{\circ}\text{F}$), producing a very dense, inert product.

An especially pure form of magnesium oxide can be made by taking the product from any of the above reactions and making a slurry with water. A slurry is a mud-like mixture of a solid and liquid that normally do not form a solution. Various chemicals can then be added to the magnesium oxide slurry to remove any contaminants, and the purified slurry is allowed to dry.

COMMON USES AND POTENTIAL HAZARDS

Each form of magnesium oxide has its specialized uses:

- Dead-burned: As refractory brick for cement kilns, furnaces, crucibles, and equipment used in the manufacture of steel;
- Hard-burned: In the production of fertilizers and animal feed, in the extraction of uranium oxide from uranium

Interesting Facts

- Magnesium oxide is sometimes used as a gemstone called periclase. It is found in a wide range of colors from colorless or white to

yellow to brown. Its use as a gemstone is somewhat limited, however, because it is not very hard.

ore, as a catalyst, in the manufacture of ceramics, for the tanning of leather, in the synthesis of magnesium compounds, and (its most important single use) in pollution control devices that remove sulfur dioxide from plant exhaust gases;

- Light-burned: In the processing of paper and pulp; as a filler in products made of rubber; and as an ingredient in a host of household and personal care products such as dusting powders, cosmetics, and pharmaceuticals. Some of the best-known pharmaceuticals containing magnesium oxide are antacids used to treat heartburn, upset stomach, or acid indigestion and laxatives for the treatment of constipation or in preparation for surgery.
- Fusion: As refractory linings for electric arc furnaces and in insulating materials used in many household electrical products.

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

REFRACTORY MATERIAL A material with a high melting point, resistant to melting,

often used to line the interior of industrial furnaces.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

Contact with magnesium oxide dust or fumes may irritate the skin, eyes, or respiratory system. Symptoms of exposure may include fatigue and lethargy. The greatest concern for such health hazards rests with people who come into contact with the pure product in their line of work.

FOR FURTHER INFORMATION

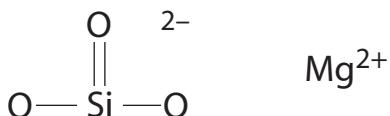
“Everything You Wanted to Know about Magnesium Oxide.” Martin Marietta Magnesia Specialties.

<http://www.magspecialties.com/students.htm> (accessed on October 14, 2005).

“Magnesia Magnesium Oxide (MgO).” Azom.com.

<http://www.azom.com/details.asp?ArticleID=54> (accessed on October 14, 2005).

See Also Magnesium Hydroxide



OTHER NAMES:
See Overview.

FORMULA:
 $\text{Mg}_3\text{Si}_4\text{O}_{10}(\text{OH})_2$

ELEMENTS:
Magnesium, silicon, oxygen, hydrogen

COMPOUND TYPE:
Hydrated salt
(inorganic)

STATE:
Solid

MOLECULAR WEIGHT:
379.27 g/mol

MELTING POINT:
1500°C (2700°F);
begins to lose water
of hydration above
900°C (1600°F)

BOILING POINT:
Not applicable

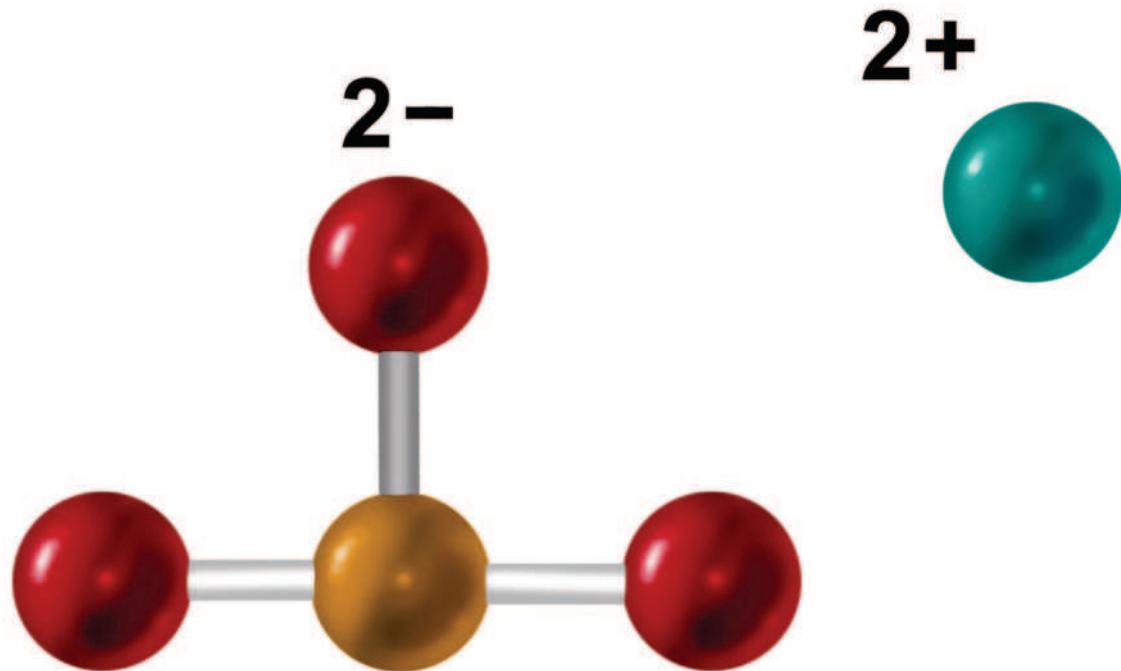
SOLUBILITY:
Insoluble in water and
most organic
solvents

KEY FACTS

OVERVIEW

Magnesium silicate hydroxide (mag-NEE-zee-um SILL-uh-kate hye-DROK-side) is also known as hydrated magnesium silicate, hydrous magnesium silicate, magnesium silicate hydrous, talc, talcum, and soapstone. It belongs to a large family of magnesium silicates that occur in nature. Magnesium silicates contain at least one magnesium ion and one or more silicate (SiO_3) ions, and often contain one or more molecules of water of hydration. Other members of the family include magnesium metasilicate (MgSiO_3), magnesium orthosilicate (Mg_2SiO_4), magnesium trisilicate ($\text{Mg}_2\text{Si}_3\text{O}_8$), and magnesium trisilicate pentahydrate ($\text{Mg}_2\text{Si}_3\text{O}_8 \cdot 5\text{H}_2\text{O}$).

The naturally-occurring form of magnesium silicate hydroxide is called talc, a soft mineral that feels waxy and soapy to the touch. This characteristic has led to another name for the mineral: soapstone. Talc's chemical formula differs somewhat from that of magnesium silicate hydroxide: $\text{Mg}_3\text{Si}_4\text{O}_{10}(\text{OH})_2$.



Magnesium silicate. Red atoms are oxygen; turquoise atom is magnesium; and orange atom is silicon. Gray stick indicates a double bond.

PUBLISHERS RESOURCE GROUP

Talc is one of the softest minerals known. It has a numerical rank of 1 on the Mohs scale of minerals. The Mohs scale ranks minerals from the softest (1 = talc) to the hardest (10 = diamond). Talc is so soft that it can be scratched with the fingernail. The mineral has a pearly luster and may come in a variety of colors, ranging from white gray, or silver to black, brown, pink, or green. Color variations depend on impurities in the mineral.

Magnesium silicate hydroxide normally occurs as a fine white powder with no odor, insoluble in most solvents, non-combustible, resistant to heat, and with a tendency to absorb moisture from the air.

HOW IT IS MADE

The magnesium silicates are obtained from natural sources, such as talc (magnesium silicate hydroxide), enstatite (magnesium metasilicate), forsterite (magnesium orthosilicate), meerschaum (magnesium trisilicate), or serpentine ($\text{Mg}_3\text{Si}_2\text{O}_7$). In all cases, extraction of the desired compound

Interesting Facts

- The magnesium silicates contain three of the eight most abundant elements in Earth's crust: oxygen (the most abundant element), silicon (the second most abundant element), and magnesium (the eighth most abundant element).

involves a number of steps in which the mineral is separated from waste products, crushed, purified, and processed into its desired form (powder or crystal, for example).

COMMON USES AND POTENTIAL HAZARDS

The form of magnesium silicate hydroxide with which most people are familiar is talc, the main ingredient in talcum powder. Talcum powder is used directly as a skin treatment, especially for babies, primarily as a moisture absorbent and for soothing the skin. It is also used in a number of cosmetic products, including blushes, eye shadows, make-up foundations, and face powders. Talc is also used widely as an additive to give products a smoother texture. Some products in which it is found include paints, rubber products, roofing materials, ceramics and insecticides. Other applications of talc include:

- As a food additive to prevent foods from clumping and sticking together;
- In the polishing of rice;
- In the production of olive oil to improve the product's clarity;
- For countertops in chemical laboratories;
- As a lubricant between sheeted products, such as particleboard, to prevent individual sheets from sticking to each other;
- As a filler in a number of products, including soap, putty, plaster, oilcloth, and rubber products; and
- As a non-caking agent in animal feeds and fertilizers.

Words to Know

ION An atom or a group of atoms with either a positive or negative charge, because it has either lost or gained an electron.

MOHS SCALE A numerical scale used to compare the hardness of a material. Talc has a Mohs value of 1; diamond has a value of 10.

REFRACTORY MATERIAL A material with a high melting point, resistant to melting, often used to line the interior of industrial furnaces.

WATER OF HYDRATION Water that has combined with a compound by some physical means.

Magnesium silicate hydroxide and other forms of magnesium silicates have a number of industrial applications, including:

- In the manufacture of glass and ceramic materials;
- As an insulating material for electrical devices;
- As a refractory material in industrial furnaces;
- In clean-up operations following oil spills; and
- As an odor absorbent.

Exposure to talc dust may cause irritation of the skin, respiratory system, and, especially, the eyes. Harmful effects may include skin rash, eye damage, and coughing and wheezing. These symptoms generally occur only as the result of long-term and consistent exposure to dust. There is no evidence that talc is carcinogenic unless contaminated with other minerals, such as asbestos and/or silica. The amount of talc and other magnesium silicates to which the average person is exposed is probably too low to produce any detectable health effects.

FOR FURTHER INFORMATION

“Chemical Summary: Talc.” CHEC’s HealtheHouse.

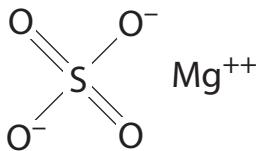
http://www.checcnet.org/healthehouse/chemicals/chemicals_detail.asp?Main_ID=2 (accessed on October 26, 2005).

“Magnesium Silicates.” In Pradyot Patnaik. *Handbook of Inorganic Chemicals*. New York: McGraw Hil, 2003, 534 535.

MAGNESIUM SILICATE HYDROXIDE

“Talc.” New Jersey Department of Health and Senior Services.
<http://www.state.nj.us/health/eoh/rtkweb/1773.pdf> (accessed on October 26, 2005).

“Talc Mineral Data.” Mineral of the Month Club.
<http://webmineral.com/data/Talc.shtml> (accessed on October 26, 2005).



OTHER NAMES:	None
FORMULA:	MgSO ₄
ELEMENTS:	Magnesium, sulfur, oxygen
COMPOUND TYPE:	Inorganic salt
STATE:	Solid
MOLECULAR WEIGHT:	120.40 g/mol
MELTING POINT:	1,127°C (2,061°F); decomposes
BOILING POINT:	Not applicable
SOLUBILITY:	Soluble in water, ethyl alcohol, and glycerol; slightly soluble in ether

KEY FACTS

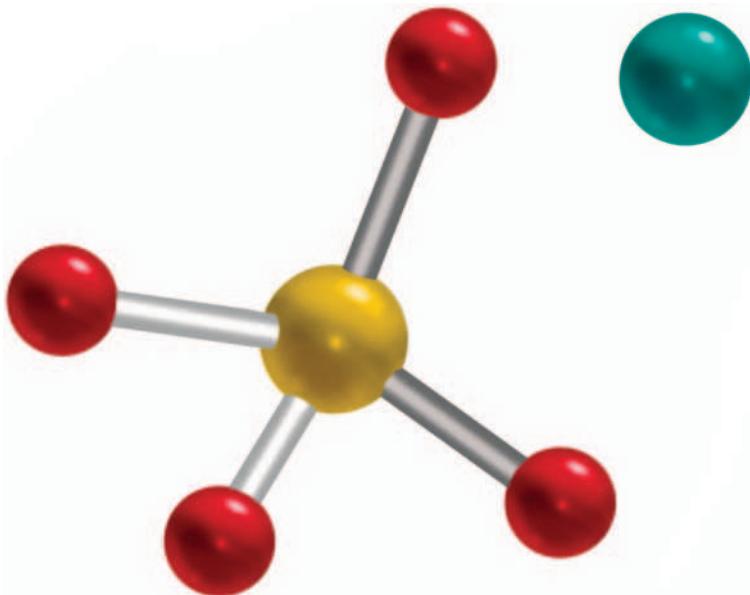
Magnesium Sulfate

OVERVIEW

Magnesium sulfate (mag-NEE-zee-um SUL-fate) occurs as the anhydrous salt and in a number of hydrated forms, including MgSO₄·H₂O, MgSO₄·4H₂O, MgSO₄·5H₂O, MgSO₄·6H₂O, and MgSO₄·7H₂O. The heptahydrate (MgSO₄·7H₂O) is commonly known as Epsom salts. All of the hydrates occur in nature as the minerals, respectively, kieserite, starkeyite, pentahydrite, hexahydrite, and epsomite. Magnesium sulfate in all forms is a colorless or white crystalline or powdery material with no odor but a bitter taste. The hydrates all lose their water of hydration when heated. For example, the heptahydrate loses one molecule of water spontaneously at room temperature, four molecules of water when heated above 70°C (160°F), six molecules of water at 150°C (300°F), and all seven molecules of water above 200°C (400°F).

Magnesium sulfate heptahydrate has been known and used since the late seventeenth century. The term *Epsom salts* was introduced in 1695 by the English naturalist Nehemiah Grew (1628–1711), who named the compound after

Magnesium sulfate heptahydrate. Red atoms are oxygen; turquoise atom is magnesium; and yellow atom is sulfur. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP



the spring waters near Epsom, England, from which it was often extracted. Grew received a royal patent for the collection, purification, and sale of the compound.

Although Epsom salts were not given their name until the late seventeenth century, their use as a therapy for many disorders was known at least two centuries earlier. People traveled to the Epsom area, as well as to other locations where the compound occurred in spring waters, just to drink the waters. Among its other benefits, Epsom salts was regarded as a purgative, a substance that helps empty the bowels (intestines).

HOW IT IS MADE

All forms of magnesium sulfate can be obtained from natural sources in a very high degree of purity, which is their primary source for commercial and industrial applications. For example, simply evaporating the water from springs like those around Epsom results in crystalline magnesium heptahydrate of sufficient purity for many uses. Water taken from salt lakes may also contain magnesium sulfate in sufficient concentration to allow its extraction by

Interesting Facts

- Up to 2 percent of sea salt may consist of magnesium sulfate.

a process of dilution and evaporation. Anhydrous magnesium sulfate is obtained from kieserite by dissolving the mineral in water at 90°C (200°F), allowing the solution to cool, and then removing the crystals that form by filtration or centrifuging.

COMMON USES AND POTENTIAL HAZARDS

Epsom salts, the heptahydrate of magnesium sulfate, has long been used for its medicinal benefits. When put into bathwater, it helps reduce inflammation and swelling, and relaxes muscles. People still visit hot springs and other types of medicinal waters where magnesium sulfate is one of the dissolved substances that contributes to one's well-being. The compound is also sold over the counter for use at home for similar purposes.

The various forms of magnesium sulfate also have a number of important commercial and industrial uses, including:

- As an animal feed, to ensure that animals receive the magnesium they require in their diets;
- In fertilizers, with the function of treating soil for magnesium deficiencies;
- In the textile industry, where it is used for weighting and sizing silk, as a mordant, for treating finished cotton fabric, and in fireproofing fabrics;
- In the production of ceramics;
- In the preparation of a specialized form of cement known as oxysulfate cement;
- In electroplating processes;

Words to Know

ANHYDROUS Form of a compound that lacks any water of hydration.

CENTRIFUGING A method of separating the components of a mixture by spinning the mixture at a high rate of speed. Centrifugal force acts differently on each component of the mixture, causing them to separate from each other.

ELECTROPLATING A process by which a thin layer of one metal is deposited on top of a second metal by passing an electric current through a solution of the first metal.

HYDRATE A chemical compound formed when one or more molecules of water is

added physically to the molecule of some other substance.

MORDANT A substance used in dyeing and printing that reacts chemically with both a dye and the material being dyed to help hold the dye permanently to the material.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

WATER OF HYDRATION Water that has combined with a compound by some physical means.

- As a catalyst for some chemical and industrial operations; and
- For the synthesis of a number of magnesium compounds, such as magnesium stearate, used in specialized types of soaps.

In dry form, magnesium sulfate can cause irritation of the skin, eyes, and respiratory system, producing symptoms such as sneezing and dryness of the mucous membranes (the soft tissues lining the breathing and digestive passages); pain, redness and tearing of the eyes; and nausea, vomiting, abdominal cramps, and diarrhea as a result of ingesting the compound. People who use the compound should use reasonable precaution in handling the material, taking care not to inhale or ingest it.

FOR FURTHER INFORMATION

“About Epsom Salt.” Epsom Salt Council.

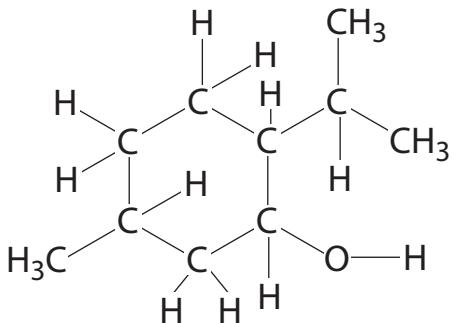
http://www.epsomsaltcouncil.org/about_epsom_salt.htm
(accessed on October 14, 2005).

"Magnesium Sulfate Anhydrous." J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/mo235.htm>
(accessed on October 14, 2005).

Nardozzi, Charlie. "Fertilize with Epsom Salts." Do It Yourself.com.

<http://doityourself.com/fertilizer/fertilizeepsomsalts.htm>
(accessed on October 14, 2005).



OTHER NAMES:

Hexahydrothymol;
methylhydroxyisopropylcyclohexane;
peppermint camphol

FORMULA:

$\text{CH}_3\text{C}_6\text{H}_9$ ($\text{C}_3\text{H}_7\text{OH}$)

ELEMENTS:

Carbon, hydrogen,
oxygen

COMPOUND TYPE:

Organic

STATE:

Solid

MOLECULAR WEIGHT:

156.26 g/mol

MELTING POINT:

41°C to 43°C (106°F to
 109°F)

BOILING POINT:

212°C (414°F)

SOLUBILITY:

Slightly soluble in
water; very soluble in
alcohol, chloroform,
and ether

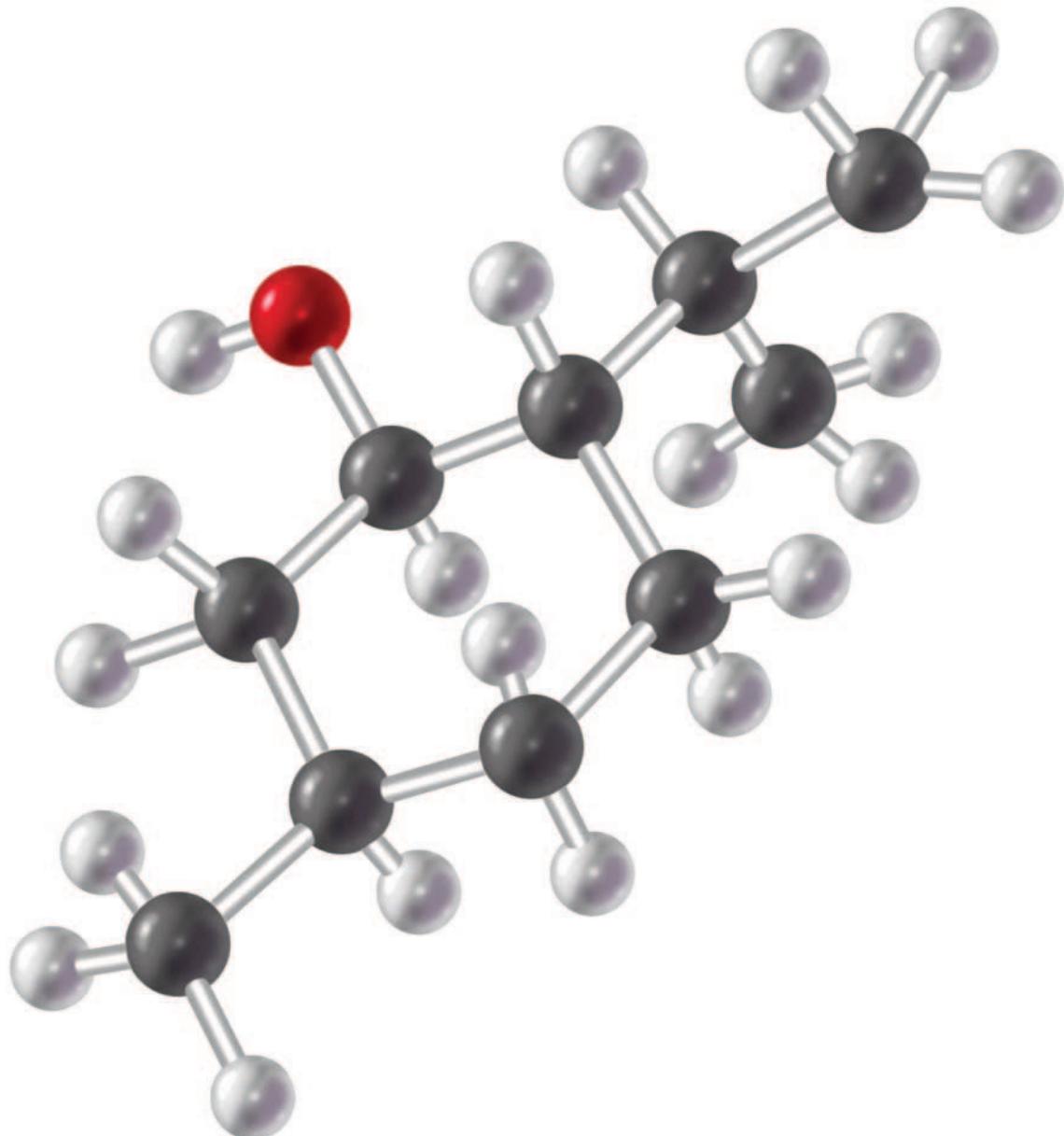
KEY FACTS

OVERVIEW

Menthol (MEN-thol) occurs naturally in the peppermint plant. In pure form it occurs as a white crystalline material with a cooling taste and odor. Peppermint is one of the oldest known herbal remedies. Dried peppermint leaves have been found in Egyptian pyramids dating to at least 1000 BCE, and its use among the Greeks and Romans in cooking and medical preparations is well known. Peppermint was not introduced to western Europe, however, until the eighteenth century, when it was used to treat a variety of ailments ranging from toothaches to morning sickness. It was first brought to the United States about a century later.

HOW IT IS MADE

Peppermint oil is extracted from the leaves of the peppermint plant, *Mentha piperita*, by steam distillation, by which various oils in the plant are separated from each other.



Menthol. Red atom is oxygen; white atoms are hydrogen; and black atoms are carbon.

PUBLISHERS RESOURCE GROUP

The peppermint oil is then frozen to extract the menthol from other components of the oil. Menthol can also be produced synthetically by the reduction of thymol $[(CH_3)_2CHC_6H_3(CH_3)OH]$ with hydrogen.

Interesting Facts

- Menthol was first added to cigarettes in the early 1900s to make a cooler tasting product. Tobacco companies claimed that menthol cigarettes were less irritating to the throat

and an effective way of treating sore throats. Today, about 25 percent of all cigarettes sold in the United States are mentholated.

COMMON USES AND POTENTIAL HAZARDS

Menthol smells like mint and creates a soothing and sometimes tingling sensation when it touches the skin. Scientists theorize that menthol creates the cooling sensation by triggering the same receptors on skin that tell the body's nerves to respond to cold temperatures.

The cooling sensation makes menthol a desirable additive to aftershave lotions, skin cleansers, lotions, sore throat lozenges, and lip balms. Menthol is also used in a variety of cosmetics applied to the skin and medications for the relief of itching. It is also added to foods such as chewing gums and candies to impart a mint-like flavor.

When inhaled or ingested as a lozenge, menthol can relieve nasal congestion and coughs, as well as cool and numb the throat to ease the pain of sore throats. It can also be used in ointments with camphor and eucalyptus to produce cooling and antiseptic properties. These ointments can be applied to the chest and/or nostrils to clear the nose and reduce coughing. One of the most famous menthol-containing products is Vicks® VapoRub, which is used to relieve coughs and congestion.

Although menthol is soothing and cooling in small quantities, it produces a quite different effect in larger quantities. Gargling with a large amount of menthol-containing mouthwash, for example, can create an unpleasant burning sensation.

Although menthol has been classified as a “generally recognized as safe” (GRAS) product and approved for use in foods by the U.S. Good and Drug Administration, some side effects have been reported. On contact with the skin,

Words to Know

DISTILLATION A process of separating two or more substances by boiling the mixture of which they are composed and condensing the vapors produced at different temperatures.

REDUCTION The chemical reaction in which oxygen is removed from

a substance or electrons are added to a substance.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

menthol may cause irritation. Ingesting large quantities can cause abdominal pain, nausea, vomiting, dizziness, drowsiness, and even coma. These effects are more likely to occur in infants and children than in adults.

FOR FURTHER INFORMATION

“Cool Menthol 1” and “Cool Menthol 2.” Great Moments in Science.
<http://www.abc.net.au/science/k2/moments/s537539.htm> and
<http://www.abc.net.au/science/k2/moments/s537548.htm>
(accessed on October 12, 2005).

Cummings, Linda Scott. “M2258 Menthol.” Paleoresearch Institute.
<http://www.paleoresearch.com/MSDS/MENTHOL%20%20SIGMA%20CHEMICAL%20%2005191997.htm> (accessed on October 12, 2005).

“Menthol and Tobacco Smoking.”
<http://goodhealth.freeservers.com/MethTobaccoIntro.html>
(accessed on October 12, 2005).

“Menthol Its Chemistry and Many Uses.”
<http://goodhealth.freeservers.com/MenthUseThisOne.htm>
(accessed on October 12, 2005).

Travis, J. “Cool Discovery: Menthol Triggers Cold Sensing Protein.”
Science News (February 16, 2002): 101–102.

See Also Camphor

OTHER NAMES:	Mercuric sulfide; cinnabar; vermillion; Chinese red
FORMULA:	HgS
ELEMENTS:	Mercury, sulfur
COMPOUND TYPE:	Inorganic salt
STATE:	Solid
MOLECULAR WEIGHT:	232.66 g/mol
MELTING POINT:	Data differ significantly; see Overview
BOILING POINT:	Not applicable
SOLUBILITY:	Insoluble in water, alcohol, and most acids; soluble in aqua regia

KEY FACTS

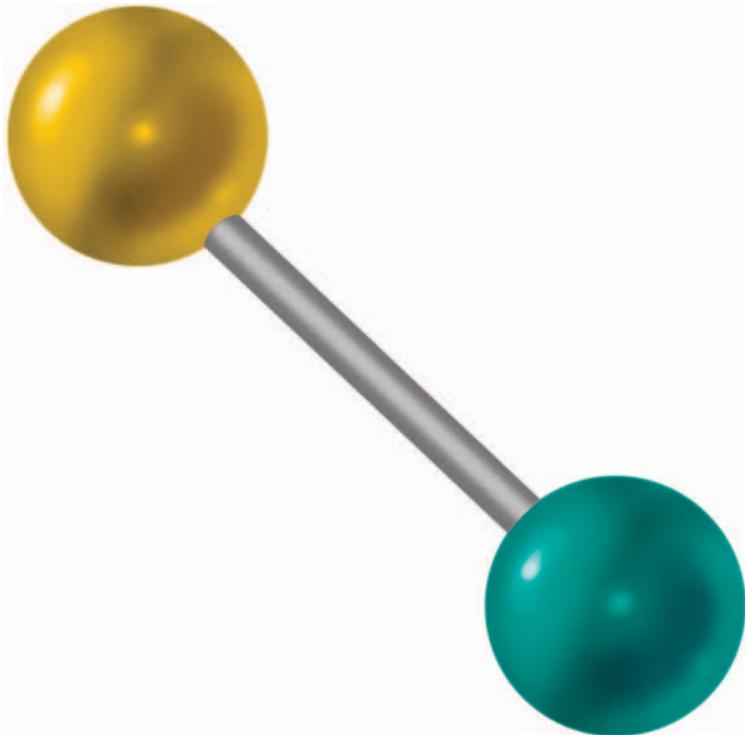
OVERVIEW

Mercury(II) sulfide (MER-kyuh-ree two SUL-fide) occurs in two forms, red and black. Red mercury(II) sulfide, commonly known as cinnabar, begins to change color when heated to temperatures of about 250°C (500°F) and converts to the black form at 386°C (727°F). If heated further, it sublimes (changes directly from a solid to a gas without first melting) at 583.5°C ($1,082^\circ\text{F}$). If allowed to cool, it then reverts to its original reddish color. Black mercury(II) sulfide goes through a similar process, changing to its red counterpart before melting at 583.5°C ($1,082^\circ\text{F}$). Some authoritative resources give significantly different temperatures for these transitions. Red mercury(II) sulfide occurs naturally as the mineral cinnabar, while black mercury(II) sulfide occurs only rarely in nature, then as the mineral metacinnabar (meaning “similar to cinnabar”).

HOW IT IS MADE

Red mercury(II) sulfide is obtained commercially from the mineral cinnabar. The compound can also be made

Mercury Sulfide. Turquoise atom is mercury and yellow atom is sulfur. Stick is a double bond. PUBLISHERS RESOURCE GROUP



synthetically by heating mercury and sulfur together in a gaseous state or by heating mercury with a solution of potassium pentasulfide (K_2S_5). The compound produced by either of these methods is commonly known as English vermillion, or simply, vermillion. The term vermillion is generally reserved for any form of mercury(II) sulfide that has been made synthetically rather than extracted from cinnabar. Other methods for the preparation of both red and black mercury(II) sulfide are available. For example, the black form can be produced by reacting sodium thiosulfate ($Na_2S_2O_3$) with sodium mercurichloride (Na_2HgCl_4).

COMMON USES AND POTENTIAL HAZARDS

The earliest records of the use of mercury(II) sulfide by humans date to about the third millennium BCE in China, where the compound was used to cure diseases, relieve pain, as a narcotic and an antiseptic, and as a preservative. Chinese alchemists referred to the compound as “celestial

Interesting Facts

- Women used vermillion during the Renaissance period to redden their lips and cheeks.
- Mercury(II) sulfide is a comparatively expensive compound, selling in late 2005 for about 1,600 dollars per 100 grams.

sands” or “god’s sand” and believed that it could transform base metals, like iron and lead, into precious metals, like silver and gold.

Today, the primary use of mercury(II) sulfide is in the production of metallic mercury. The sulfide is heated in a furnace to temperatures of 600°C to 700°C (1,100°F to 1,300°F), resulting in the formation of sulfur dioxide and mercury metal. In a second process, the sulfide is treated with lime (CaO), resulting in the formation of mercury metal, calcium sulfide (CaS) and calcium sulfate (CaSO₄).

The other major use for mercury(II) sulfide, in either red or black form, is as a pigment in artists’ paints, for coloring paper and plastics, and for marking linen. The black form is also used as a pigment for the coloring of rubber, horn, and other materials. Red mercury(II) sulfide finds some use also as an antibacterial agent.

Both forms of mercury(II) sulfide are highly toxic by ingestion, inhalation, or absorption through the skin. Some of the symptoms resulting from mercury(II) sulfide poisoning include inflammation and itching of the skin; redness,

Words to Know

ALCHEMY An ancient field of study from which the modern science of chemistry evolved.

AQUA REGIA A combination of concentrated nitric acid and hydrochloric acid.

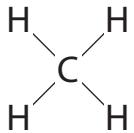
itching, burning, and watering of the eyes; excessive salivation; pain when chewing; gingivitis with loosening of teeth; and mental disturbances, such as loss of memory, insomnia, irritability, and vague fears of depression. Anyone who has been exposed to mercury(II) sulfide and experiences such symptoms requires immediate medical attention.

FOR FURTHER INFORMATION

Liu, Guanghua. "Chinese Cinnabar." *The Mineralogical Record* (January February 2005): 69 80.

"The Mineral Cinnabar." Amethyst Galleries.

<http://mineral.galleries.com/minerals/sulfides/cinnabar/cinnabar.htm> (accessed on October 14, 2005).



OTHER NAMES:	Marsh gas; methyl hydride
FORMULA:	CH_4
ELEMENTS:	Carbon, hydrogen
COMPOUND TYPE:	Hydrocarbon; alkane (organic)
STATE:	Gas
MOLECULAR WEIGHT:	16.04 g/mol
MELTING POINT:	-182.47°C (-296.45°F)
BOILING POINT:	-161.48°C (-322.63°F)
SOLUBILITY:	Very slightly soluble in water and acetone; soluble in ethyl alcohol, methyl alcohol, and ether

K E Y F A C T S

Methane

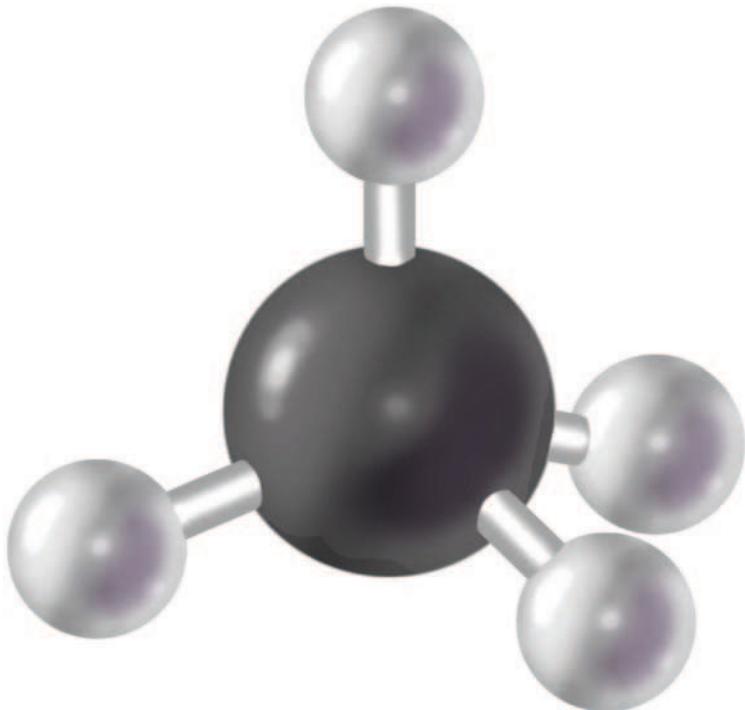
OVERVIEW

Methane (METH-ane) is a colorless, odorless, tasteless, flammable gas that is less dense than air. It is the primary component of natural gas. Methane is the simplest of all hydrocarbons, organic compounds that contain carbon and hydrogen and no other elements.

HOW IT IS MADE

Methane formed millions of years ago from microscopic underwater plants and bacteria that dropped to the bottom of the ocean when they died. Over millions of years, they were crushed and heated by the pressure of layers of sand, dirt, and other materials that accumulated on top of them. The mineral components of the undersea mud gradually turned into a type of rock known as shale. Some of the organic components turned into natural gas, which is mostly methane. The natural gas became trapped in porous rocks called reservoir rocks and in larger pockets of the rock called reservoirs or geologic traps. Natural gas is now found in

Methane. Black atom is carbon and white atoms are hydrogen. PUBLISHERS RESOURCE GROUP



pockets by itself, but is more commonly found floating on top of petroleum lakes in underground reservoirs.

Methane is also found in conjunction with pockets of coal. The largest reserves of natural gas in the United States are in Texas, Alaska, Oklahoma, Ohio, and Pennsylvania. Oil and gas companies remove natural gas from the ground by drilling. They then purify the natural gas by separating the components of which it is made, such as methane, ethane, propane, and butane. After isolation from natural gas, methane is often liquefied, which makes its easier to store and transport.

Although abundant supplies of methane exist, it can also be produced synthetically. For example, the reaction between steam and hot coal results in the formation of synthesis gas, a mixture of hydrogen and carbon monoxide. When this mixture is passed over a catalyst containing nickel metal, methane is formed. A very similar process, called the Sabatier process, uses a mixture of hydrogen and carbon dioxide, rather than carbon monoxide, also resulting in the formation

Interesting Facts

- Methane is sometimes called marsh gas because it forms in swamps as plants and animals decay under water.
- Methane is odorless, but gas companies add traces of sulfur-containing compounds with strong odors so that people will be able to smell gas leaks and avoid suffocation or explosions.
- Some experts estimate that enough methane is present in the Earth's surface to last as much as two hundred years, although extraction of some methane resources may prove to be difficult.

of methane. Finally, methane produced during the anaerobic decomposition of manure can be captured and purified.

COMMON USES AND POTENTIAL HAZARDS

When methane burns, it releases a large amount of energy, making it useful as a fuel. Humans have known about methane as a source of energy for thousands of years. Temples in the ancient world often burned "eternal flames" that may have been fueled by natural gas. In the early nineteenth century, people began using natural gas as a light source. Once oil was discovered in the 1860s, however, its use, and the electricity produced by burning oil, became much more popular, and people abandoned natural gas as a fuel except for limited use in cooking.

Natural gas has become more popular in recent years because its use results in less pollution than petroleum and other fossil fuels. Some uses include heating homes, offices, and factories; powering room heaters and air conditioners; and operating home appliances such as water heaters and stoves.

Scientists are now exploring other uses for methane and natural gas with the hope that they might eventually become the most important fuels used by humans. Methane has some advantages over petroleum and coal as a fuel. It

Words to Know

ANAEROBIC Describing a process that takes place in the absence of oxygen.

GREENHOUSE GAS One of several gases, including carbon dioxide and ozone, that causes the greenhouse effect on Earth.

burns more cleanly than either of these other fossil fuels, producing only carbon dioxide and water as combustion products. Some experts believe that methane could be used as a power source of fuel cells, cells that burn hydrogen to produce electricity. Adding natural gas to oil- or coal-fired burners would also help reduce the greenhouse gas emissions of these appliances.

In addition to its applications as a fuel, methane is used in the manufacture of a number of organic and inorganic compounds. For example, ammonia, which is the tenth most important chemical compounds in the United States, based on quantity produced, is made from hydrogen and nitrogen gases. Over 90 percent of the hydrogen used to make ammonia is now obtained by reacting methane with water at high temperatures over a catalyst of iron oxide (Fe_3O_4). Other compounds produced from methane include methanol (methyl alcohol), acetylene (ethyne), formaldehyde (methanal), hydrogen cyanide, carbon tetrachloride, chloroform, methylene chloride, and methyl chloride.

Methane is not toxic, but it can cause suffocation by reducing or eliminating the oxygen a person needs to breathe normally. The primary hazard posed by the gas is its flammability and explosive tendency.

FOR FURTHER INFORMATION

“Chemical of the Week: Methane.” Science Is Fun.

<http://scifun.chem.wisc.edu/chemweek/methane/methane.html> (accessed on October 17, 2005).

“Methane.” U.S. Environmental Protection Agency.

<http://www.epa.gov/methane/> (accessed on October 17, 2005).

"Methane Madness: A Natural Gas Primer." The Coming Global Oil Crisis.
<http://www.oilcrisis.com/gas/primer/> (accessed on October 17, 2005).

Sherman, Joseph, and Steve Brick. *Fossil Fuel Power*. Mankato, MN: Capstone Press, 2003.

See Also Propane

**OTHER NAMES:**

Methanol; wood alcohol; wood spirit; carbinol

FORMULA:

CH_3OH

ELEMENTS:

Carbon, hydrogen, oxygen

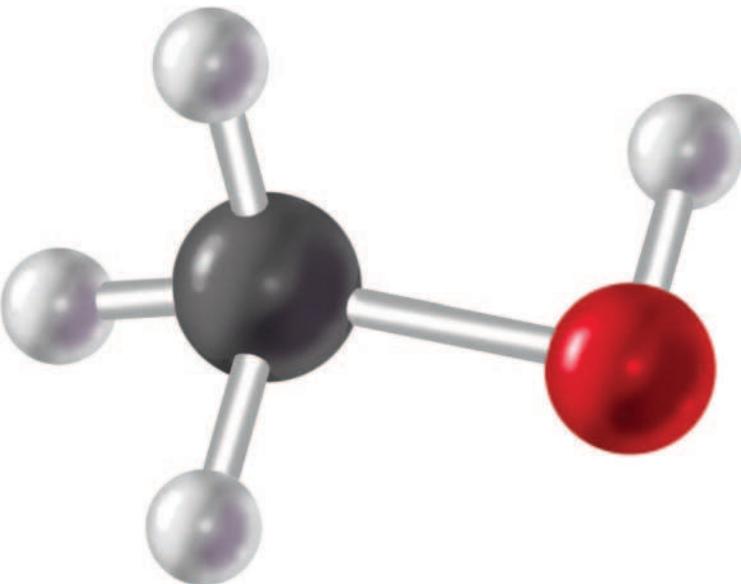
COMPOUND TYPE:
Alcohol (organic)**STATE:**
Liquid**MOLECULAR WEIGHT:**
32.04 g/mol**MELTING POINT:**
 -97.53°C (-143.6°F)**BOILING POINT:**
 64.6°C (148°F)**SOLUBILITY:**
Miscible with water, ethyl alcohol, ether, acetone, and many other organic solvents**K E Y F A C T S****OVERVIEW**

Methyl alcohol (METH-uhl AL-ko-hol) is a clear, colorless, flammable, toxic liquid with a slightly alcoholic odor and taste. Methyl alcohol is the simplest alcohol, a family of organic compounds characterized by the presence of one or more hydroxyl (-OH) groups.

HOW IT IS MADE

Methyl alcohol occurs naturally in plants and animals, including humans, as the product of metabolic reactions that occur in all organisms. It also occurs in the atmosphere as the result of the decomposition of dead organisms in the soil. None of these sources is utilized for the commercial production of methyl alcohol. Instead, the primary method for the preparation of methyl alcohol is to react carbon monoxide with water at a temperature of about 250°C (480°F) and pressures of 50 to 100 atmospheres over a mixed catalyst of copper, zinc oxide, and aluminum oxide. Efforts are being made to develop other methods of synthesizing methyl alcohol. In one process, for

Methyl alcohol. Red atom is oxygen; white atoms are hydrogen; and black atom is carbon. PUBLISHERS RESOURCE GROUP



example, simple hydrocarbons, such as methane, are oxidized over a catalyst of molybdenum metal to produce the alcohol. None of the experimental methods developed for the production of methyl alcohol can yet compete with the traditional carbon monoxide-hydrogen process, however.

COMMON USES AND POTENTIAL HAZARDS

Consumption of methyl alcohol in the United States for 2005 reached about 12 billion liters (3 billion gallons). The largest demand for the compound was in the production of MTBE (methyl-*tert*-butyl ether), a gasoline additive used to improve the efficiency with which a fuel burns and to reduce pollutants released to the atmosphere. Demand for the additive increased rapidly after the U.S. Congress passed the 1990 Clean Air Act Amendments requiring significant reductions in the release of certain pollutants into the atmosphere. A decade later, however, serious questions were being raised about possible serious environment dangers posed by MTBE released into the soil. In the last few years, enthusiasm for the use of MTBE has begun to disappear and a number of states have adopted bans on its use as a gasoline additive. As a result of these actions, the demand for methyl alcohol in producing MTBE has dropped dramatically in the last few years.

Interesting Facts

- At one time, the primary method for making methyl alcohol was to heat wood in a closed space, accounting for the compound's common and popular name of "wood alcohol."
- Methyl alcohol was first isolated, although not in a pure form, by the English chemist and physicist Robert Boyle (1627–1691) although the compound was not synthesized for another two centuries. It was then produced by the French chemist Pierre Eugène Marcellin Berthelot (1827–1907).

The next most important demand for methyl alcohol is as a raw material in the synthesis of many important organic compounds, including formaldehyde; acetic acid; chloromethanes, compounds in which the hydroxyl group and/or one or more hydrogen has been replaced by fluorine, chlorine, bromine, and/or iodine; methyl methacrylate, a compound from which acrylic plastics are made; methylamines, the source of another important class of plastics, dimethyl terephthalate, the monomer for yet another class of plastics; and other products.

Relatively small quantities of methyl alcohol are used in a number of other applications, including:

- As a solvent for household and industrial products;
- As a deicing agent;
- In the preparation of embalming fluids;
- As a softening agent for plastics;
- As a fuel for camp stoves, soldering torches, and race cars;
- In paint removing products;
- As an antifreeze and windshield washing fluid; and
- In the manufacture of a number of pharmaceuticals, including streptomycin, vitamins, and hormones.

Words to Know

METABOLISM A biological process that includes all of the chemical reactions that occur in cells by which fats, carbohydrates, and other compounds are broken down to produce energy and the compounds needed to build new cells and tissues.

MISCIBLE able to be mixed; especially applies to the mixing of one liquid with another.

MONOMER A small molecule used in polymerization reactions to produce

very large molecules in which the monomer is repeated hundreds or thousands of times.

SOLVENT A liquid that dissolves another substance.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

Methyl alcohol poses both safety and health hazards. It is highly flammable and, with the appropriate mixture of air, explosive. It is also very toxic by ingestion, producing a variety of effects that include blurred vision, headache, dizziness, drowsiness, and nausea. Although a fatal dose is usually in the range of 100 to 250 mL, cases have been reported in which a person has died after consuming less than 30 mL of the compound. People who work with methyl alcohol in their jobs, including bookbinders, dyers, foundry workers, gilders, hat makers, ink makers, laboratory technicians, painters, photoengravers, and chemical workers, are especially at risk from methanol poisoning. The ready availability of the compound and products in which it is an ingredient means that everyone who uses such products should be aware of the health risks involved in its use. Medical attention is required immediately in case of the ingestion of methyl alcohol.

FOR FURTHER INFORMATION

"Material Safety Data Sheet: Methyl Alcohol, Reagent ACS, 99.8% (GC)." Department of Chemistry, Iowa State University.
<http://avogadro.chem.iastate.edu/MSDS/methanol.htm>
(accessed on October 17, 2005).

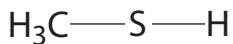
McGrath, Kimberley A. "Methanol." *World of Scientific Discovery*, 2nd edition. Detroit, MI: Gale, 1999.

“Methanol.” U.S. Environmental Protection Agency. Technology Transfer Network, Air Toxics Website.
<http://www.epa.gov/ttn/atw/hlthef/methanol.html> (accessed on October 17, 2005).

Salocks, Charles, and Karlyn Black Kaley. “Methanol.” Technical Support Document: Toxicology Clandestine Drug Labs/Methamphetamine, Volume 1, Number 10.
http://www.oehha.ca.gov/public_info/pdf/TSD%20Methanol%20Meth%20Labs%2010'8'03.pdf (accessed on October 17, 2005).

U. S. Department of Health and Human Services. “Methanol Toxicity.” *American Family Physician* (January 1993): 163–171.

See Also Carbon Monoxide; Formaldehyde

**OTHER NAMES:**

Methanethiol;
mercaptomethane;
thiomethyl alcohol;
methyl sulfhydrate

FORMULA:

CH_3SH

ELEMENTS:

Carbon, hydrogen,
sulfur

COMPOUND TYPE:

Mercaptan; thiol
(organic)

STATE:

Gas

MOLECULAR WEIGHT:

48.11 g/mol

MELTING POINT:

-123°C (-189°F)

BOILING POINT:

5.9°C (43°F)

SOLUBILITY:

Slightly soluble in
water; soluble in
ethyl alcohol and
ether

**K
E
Y
F
A
C
T
S****OVERVIEW**

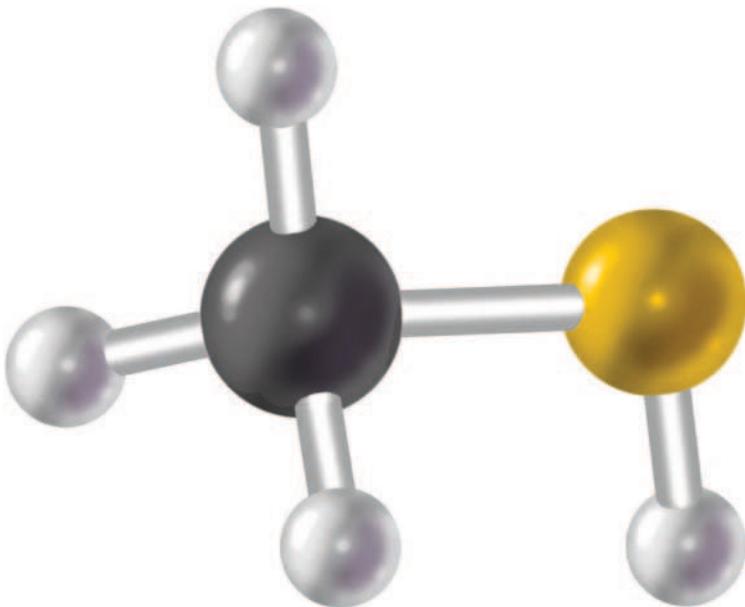
Methyl mercaptan (METH-uhl mer-KAP-tan) is a colorless, highly flammable, foul-smelling gas with the odor of rotten cabbage released from decaying animal and vegetable matter. It is also produced in the intestinal tract by the action of bacteria on a variety of proteins known as the albumins.

Methyl mercaptan belongs to a class of organic compounds called mercaptans or thiols in which one or more sulphydryl (-SH) groups are attached to a carbon atom. Methyl mercaptan has only one carbon atom, but some mercaptans contain up to twenty carbon atoms. Like methyl mercaptan, other mercaptan compounds are known for their disagreeable odors. For example, allyl mercaptan has the characteristic smell of garlic, while butyl mercaptan occurs in the spray that skunks release to protect themselves from predators.

HOW IT IS MADE

Methyl mercaptan is made by the direct reaction between methanol (methyl alcohol; CH_3OH) and hydrogen sulfide gas

Methyl mercaptan. White atoms are hydrogen; black atom is carbon; and yellow atom is sulfur. PUBLISHERS RESOURCE GROUP



(H₂S). The hydroxyl group (-OH) from the alcohol combines with one hydrogen from hydrogen sulfide to form water, leaving the methyl mercaptan behind as the major product.

COMMON USES AND POTENTIAL HAZARDS

Methyl mercaptan's primary use is in the synthesis of other organic compounds, especially pesticides, fungicides, jet fuel components, plastics, and the amino acid methionine [CH₃SCH₂CH₂CG(NH₂)COOH]. Amino acids are the building-block compounds from which proteins are made. The compound is also used as an odorant, a substance with a noticeable and usually offensive odor added to odorless compounds for the purpose of safety. For example, propane and natural gas are two widely used gases that are very flammable, but odorless. If these gases were used without having an odorant added, consumers might not be aware of a leak until the gas caught fire or exploded. The presence of the odorant, such as methyl mercaptan, makes a leak obvious and allows it to be repaired before an accident occurs.

In spite of the fact that it is a natural product, methyl mercaptan is a health and environmental risk. When inhaled, it can irritate the eyes, skin, nose, throat, and lungs, producing

Interesting Facts

- Methyl mercaptan may be responsible for some of the unpleasant odors produced by humans, such as bad breath, flatulence, and stinky feet. When bacteria attack proteins in the body, they release several gases, methyl mercaptan among them. These gases are responsible for bad breath and periodontal disease (inflammation of the gums) in the mouth, and unpleasant odors from other parts of the body. For example, old or dirty socks and shoes are ideal breeding sites for the bacteria responsible for the production of methyl mercaptan and similar foul-smelling organic compounds. One popular product designed to deal with this problem is shoe insoles that contain activated charcoal. Activated charcoal is a form of charcoal consisting of very fine grains that absorbs large volumes of gases, such as methyl mercaptan and its bad-smelling cousins.
- Why does eating asparagus produce bad-smelling urine? Asparagus contains the amino acid methionine, which is metabolized in the body to produce methyl mercaptan. The peculiar odor of urine produced after consuming asparagus is caused by the methyl mercaptan excreted from the body in urine.

dizziness, headache, vomiting, muscle weakness, and loss of coordination. In large concentrations, methyl mercaptan can damage the central nervous system, leading to respiratory failure and even death. The U.S. Occupational Safety and Health Administration (OSHA) has set maximum exposure limits at 20 milligrams of the compound per cubic meter of air per eight-hour work day.

FOR FURTHER INFORMATION

"Bad Breath: What Usually Is the Source of a Person's Bad Breath?" Animated Teeth.com.
http://www.animated-teeth.com/bad_breath/t3_causes_of_halitosis.htm (accessed on October 17, 2005).

METHYL MERCAPTAN

Emsley, John. "Molecules at an Exhibition."

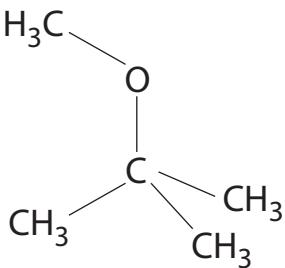
<http://www.nytimes.com/books/first/e/emsley/molecules.html>
(accessed on October 17, 2005).

"Safety Data Sheet: Methyl Mercaptan." Air Liquide.

http://www.airliquide.com/safety/msds/en/o83_AL_EN.pdf
(accessed on October 17, 2005).

"ToxFAQs for Methyl Mercaptan." Agency for Toxic Substances and Disease Registry.

<http://www.atsdr.cdc.gov/tfacts139.html> (accessed on October 17, 2005).



OTHER NAMES:

MTBE; see Overview for additional names

FORMULA:
 $(\text{CH}_3)_3\text{COCH}_3$

ELEMENTS:
Carbon, hydrogen, oxygen

COMPOUND TYPE:
Ether

STATE:
Liquid

MOLECULAR WEIGHT:
88.15 g/mol

MELTING POINT:
−108.6°C (−163.5°F)

BOILING POINT:
55.0°C (131°F)

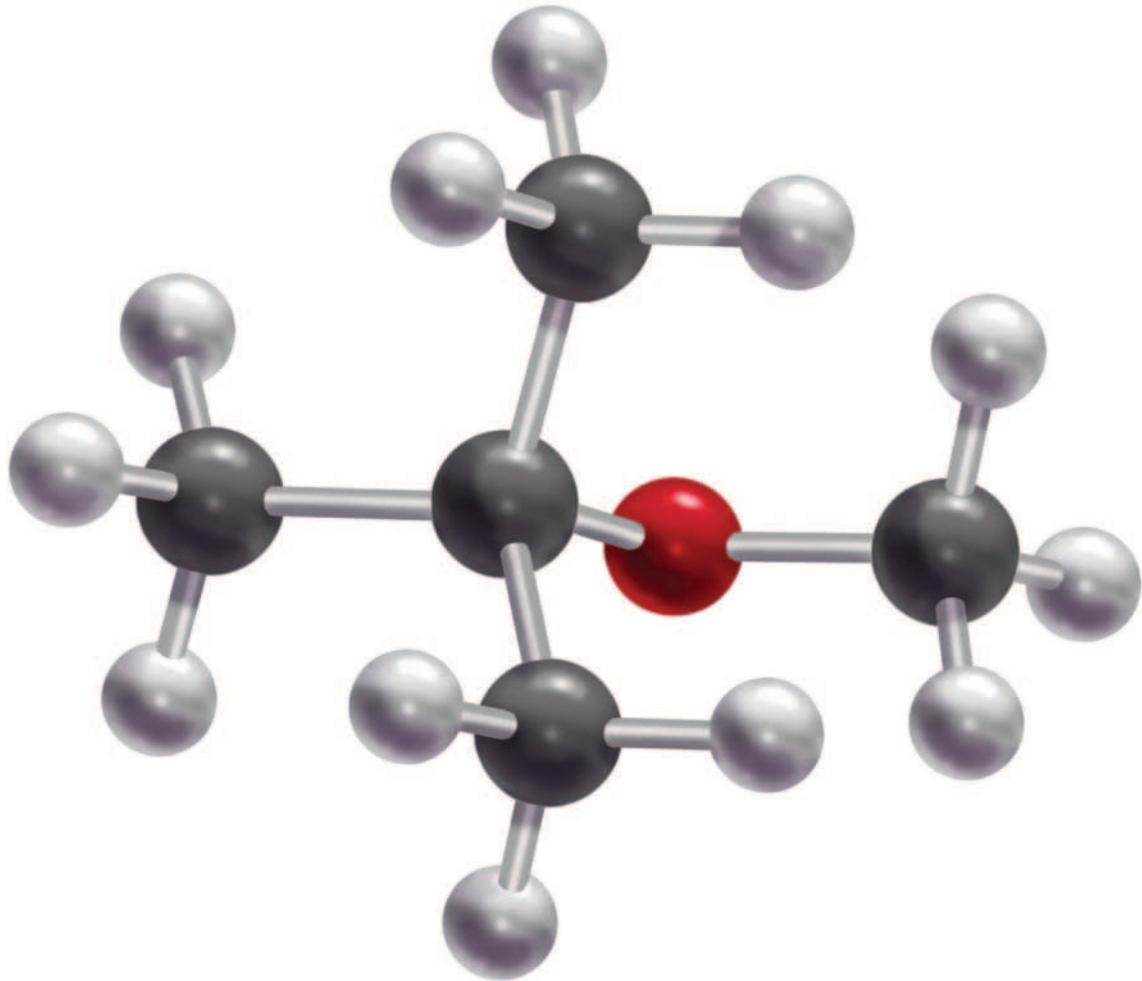
SOLUBILITY:
Soluble in water; very soluble in ethyl alcohol and ether

**K
E
Y
F
A
C
T
S**

OVERVIEW

Methyl-t-butyl ether (METH-el TER-she-air-ee BYOO-till EE-thur) is a volatile (evaporates easily), colorless, flammable liquid that forms an azeotropic mixture with water. Azeotropic mixtures are combinations of two or more liquids that boil at the same temperature and, therefore, cannot be easily separated from each other.

MTBE was first synthesized in the 1960s by researchers at the Atlantic Richfield Corporation (now ARCO) as an additive designed to increase the octane number (fuel efficiency) of gasoline. The compound was created as a replacement for tetraethyl lead ($\text{Pb}(\text{C}_2\text{H}_5)_4$), which had long been added to gasolines to improve their octane number. Tetraethyl lead was commonly called simply “lead.” From 1973 until 1996, lead was gradually removed from most gasoline because of its dangerous environmental and health effects. In 1990, a new use for MTBE was found. In that year, legislation passed by the U.S. Congress required that changes be made in the composition of gasoline so that it would burn more cleanly

**Methyl t butyl Ether (MTBE).**

Black atoms are carbon; red atom is oxygen; white atoms are hydrogen. All sticks are single bonds.

PUBLISHERS RESOURCE GROUP

and release fewer pollutants into the atmosphere. One way for companies to meet this regulation was to add oxygenates to their gasoline. Oxygenates are chemical compounds that contain oxygen. They react with fuels, giving off their oxygen and increasing the efficiency with which the fuels burn.

MTBE rapidly became a very popular chemical in the automotive fuel industry. Production rose at a rate of about 7 percent per year in the 1980s. It increased from about 500 million kilograms (1 billion pounds) in 1983 to about 3 billion kilograms (6 billion pounds) in 1990. In that year, it ranked twenty-fourth among all chemicals produced in the United States.

Interesting Facts

- Virtually all of the MTBE produced is used as a fuel additive. If proposed bans and limitations go into effect, there will no longer be any demand for the chemical. As such, production is likely to drop nearly

to zero. One company that surveys chemical production predicts that the production of MTBE is likely to decrease at the rate of about 8 percent per year over the foreseeable future.

In the early 1990s, however, MTBE began showing up in groundwater in a number of states. Upon investigation, researchers discovered that MTBE had leaked out of underground tankers, had been spilled during transportation, and had escaped into the environment in other ways. It sank into the ground, where it mixed with groundwater forming azeotropic mixtures that could not be easily separated. The compound was found in wells, lakes, streams, and public water supplies.

These findings raised concerns because some experts believe that MTBE is a health hazard to humans and other animals. Although strong evidence is not yet available, many authorities believe that MTBE is a carcinogen and that it may be responsible for other health problems and damage to the environment. These concerns have led to legislative and administrative action banning or limiting the use of MTBE in gasoline. The first such action occurred in 1999 when Governor Gray Davis of California announced a program to cut back and eventually eliminate the use of MTBE in the state. A year later, the U.S. Environmental Protection Agency announced similar plans for the nation as a whole.

MTBE is also known by the following names: Methyl-*tert*-butyl ether; *t*-butyl-methyl ether; and *tert*-butyl-methyl ether.

HOW IT IS MADE

MTBE is made by reacting methanol (methyl alcohol; CH₃OH) with isobutylene (isobutene; CH₃C(CH₃)=CH₃). The

Words to Know

AZEOTROPIC MIXTURE A combination of two or more liquids that boil at the same temperature and, therefore, cannot be easily separated from each other.

CARCINOGEN A chemical that causes cancer in humans or other animals.

OXYGENATES Chemical compounds that contain oxygen.

SOLVENT A substance that is able to dissolve one or more other substances.

VOLATILE Able to turn to vapor easily at a relatively low temperature.

formula for this reaction is $\text{CH}_3\text{OH} + \text{CH}_3\text{C}(\text{CH}_3)=\text{CH}_2 \rightarrow (\text{CH}_3)_3\text{COCH}_3$.

COMMON USES AND POTENTIAL HAZARDS

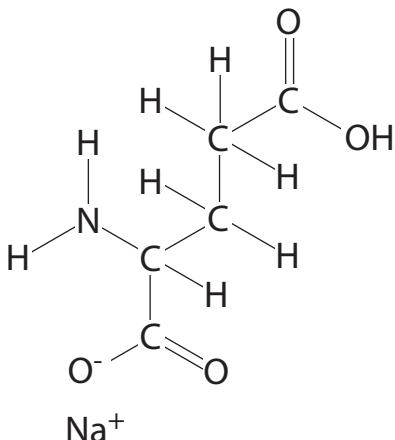
More than 99 percent of the MTBE produced is used as a gasoline additive. The remaining quantity is used as a solvent and in the production of other chemical compounds.

MTBE is a mild irritant to the skin, eyes, and respiratory system. In larger doses, it may cause more serious problems, including nausea, vomiting, diarrhea, gastrointestinal problems, headache, dizziness, and loss of balance and coordination. In extreme cases, MTBE may produce severe lung damage, respiratory failure, convulsions, respiratory arrest (breathing stops), unconsciousness, coma, and death. The compound has been found to cause cancer in laboratory animals. No similar studies are available for humans, but experts tend to agree that MTBE is likely carcinogenic in humans also.

FOR FURTHER INFORMATION

"Groundwater Protection." American Petroleum Institute.
[http://api.ep.api.org/environment/index.cfm?
 objectid=9EC44CD5 E167 49C4 8EBE7
 B354E4B3CD9&method=display_body&er=1&bitmask=
 002008008000000000](http://api.ep.api.org/environment/index.cfm?objectid=9EC44CD5E16749C48EBE7B354E4B3CD9&method=display_body&er=1&bitmask=002008008000000000) (accessed on December 29, 2005).

- "Methyl tert Butyl Ether (MTBE) and Other Gasoline Oxygenates." U.S. Geological Survey.
<http://sd.water.usgs.gov/nawqa/vocns/mtbe.html> (accessed on December 29, 2005).
- "Methyl Tertiary Butyl Ether (MTBE)." U.S. Environmental Protection Agency.
<http://www.epa.gov/mtbe/> (accessed on December 29, 2005).
- "MTBE (methyl tert butyl ether)." The Innovation Group.
<http://www.theinnovationgroup.com/welcome.htm> (accessed on December 29, 2005).



OTHER NAMES:

Sodium glutamate; glutamic acid mono-sodium salt; MSG

FORMULA:

$\text{COOH}(\text{CH}_2)_2\text{CH}(\text{NH}_2)\text{-COONa}$

ELEMENTS:

Carbon, hydrogen, oxygen, nitrogen, sodium

COMPOUND TYPE:

Organic salt

STATE:

Solid

MOLECULAR WEIGHT:

169.11 g/mol

MELTING POINT:

Not applicable; decomposes when heated

BOILING POINT:

Not applicable

SOLUBILITY:

Very soluble in water and ethyl alcohol

K
E
Y
F
A
C
T
S

OVERVIEW

Monosodium glutamate (mon-oh-SOH-dee-yum GLOO-tuh-mate) is an almost completely odorless white crystalline powder. It is the sodium salt of a common amino acid called glutamic acid. An organic salt is a compound formed when an inorganic base, such as sodium hydroxide, reacts with an organic acid, such as glutamic acid.

Monosodium glutamate has been available as a commercial product for about a century. But the compound has been used in its natural form for much longer. The ancient Greeks and Romans used fish sauce, which contains glutamic acid as a natural ingredient, in their cooking. Later Europeans also used a form of the substance in a product known as garum.

The German chemist Karl Heinrich Ritthausen (1826-1912) first identified glutamic acid in wheat gluten in 1866, and its chemical structure was first identified in 1890 by the German chemist Wolff. Then in 1908, the Japanese chemist Kikunae Ikeda (1864-1936) discovered the flavor-enhancing properties of glutamic acid. He found in seaweed broth a

MONOSODIUM GLUTAMATE



Monosodium glutamate. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; blue atom is nitrogen; and turquoise atom is sodium.

Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

compound, later identified as glutamic acid, responsible for the unique taste of cheese, meat, and tomatoes. The compound was not sweet, salty, sour, or bitter, but instead had a rich meaty taste. He named the unusual flavor *umami*. Ikeda used glutamic acid crystals to make a seasoning, which he then patented. Like salt and sugar, it was readily soluble in water and could be stored for long periods of time without clumping. The seasoning was actually the sodium salt of glutamic acid, monosodium glutamate, first sold commercially in Japan under the name *Ajinomoto*, which means “essence of taste.” The product was introduced into the United States shortly after World War II, and the U.S. Food and Drug Administration approved its use as a food additive in 1958.

HOW IT IS MADE

Glutamic acid in the form of glutamate exists in two forms: bound and free. The term bound glutamate refers to the glutamic acid that has combined with other amino acids to form proteins. The term free glutamate refers to the acid radical $\text{COOH}(\text{CH}_2)_2\text{CH}(\text{NH}_2)\text{COO}^-$ formed when a single molecule of glutamic acid loses one hydrogen ion. Foods such as

Interesting Facts

- The flavor-enhancing qualities of monosodium glutamate are a function of its concentration in foods. At concentrations of less than about 0.5 percent, it enhances the flavor of a food product. At greater concentrations, it begins to make the food product taste bad.

tomatoes, mushrooms, and cheeses are naturally high in glutamate, which is responsible for their strong flavors. Only the free form of glutamate enhances food flavors. In the human body, glutamate is a nonessential amino acid found in the brain, muscles, kidneys, liver, and other organs.

Monosodium glutamate is made commercially by the hydrolysis of the waste products of beet sugar refining, wheat, or corn gluten. Hydrolysis is the process by which a compound reacts with water to form two new compounds. In the preparation of monosodium glutamate, water breaks apart the proteins present in the raw materials used (such as wheat or corn gluten), freeing the amino acids of which they are made. The glutamic acid present in the proteins can then be separated from other amino acids present, then isolated and purified. Monosodium glutamate can also be made by a very similar process in which bacterial fermentation is used to break proteins down into their component parts, of which glutamic acid is one, a process which is now the primary means of production for the compound.

COMMON USES AND POTENTIAL HAZARDS

The sole use of monosodium glutamate is as a flavor enhancer in food products. Annual worldwide consumption of the additive has been estimated at about 1 million metric tons (1.1 million short tons). It is used primarily in a variety of Asian foods, including soups, canned foods, and processed meats.

Monosodium glutamate is considered safe for human consumption. It is on the U.S. Food and Drug Administration's Generally Recognized as Safe (GRAS) list. The GRAS

list contains chemicals that have never been tested scientifically for safety, but are generally believed to be safe for human consumption. In 1987, health experts from the United Nations and the World Health Organization reviewed more than two hundred studies on MSG and determined that the compound is a safe food additive when used at customary levels.

In spite of these studies and rulings, many people believe that monosodium glutamate can have harmful health effects. They say that eating foods that include MSG can cause headache, nausea, sweating, rapid heartbeat, a burning sensation in the back of the neck, and/or tightness in the chest. Thus far, scientific studies have been unable to confirm the association of MSG with these symptoms.

Researchers theorize that the reactions may be the result of allergies to certain ingredients in the foods. In particular, they are likely to occur among people who have specific allergies to monosodium glutamate or who have asthma.

FOR FURTHER INFORMATION

“The Facts on Monosodium Glutamate.” European Food Information Council.

<http://www.eufic.org/gb/food/pag/food35/food352.htm>
(accessed on October 17, 2005).

“FDA and Monosodium Glutamate (MSG).” U.S. Food and Drug Administration.

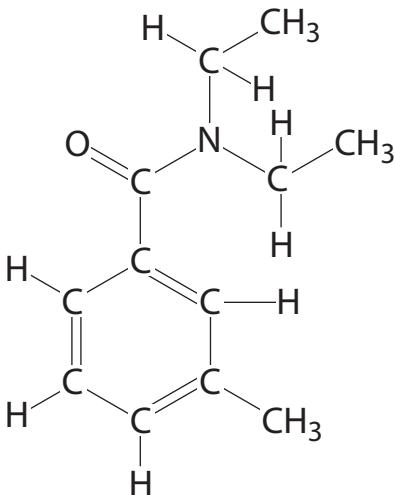
<http://www.cfsan.fda.gov/~lrd/msg.html> (accessed on October 17, 2005).

“Glutamate Facts, Information and On Line Services.” International Glutamate Information Service.

[http://www\(glutamate.org](http://www(glutamate.org) (accessed on October 17, 2005).

“L glutamic Acid, Sodium Salt.” J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/g3975.htm>
(accessed on October 17, 2005).



OTHER NAMES:

N,N-diethyl-m-toluamide; DEET

FORMULA:

C₆H₄(CH₃)CON(C₂H₅)₂

ELEMENTS:

Carbon, hydrogen, nitrogen, oxygen

COMPOUND TYPE:

Organic

STATE:

Liquid

MOLECULAR WEIGHT:

191.27 g/mol

MELTING POINT:

-45°C (-49°F)

BOILING POINT:

160°C (320°F) at 19 mm Hg pressure

SOLUBILITY:

Insoluble in water; soluble in alcohol, ether, benzene, and most other organic solvents

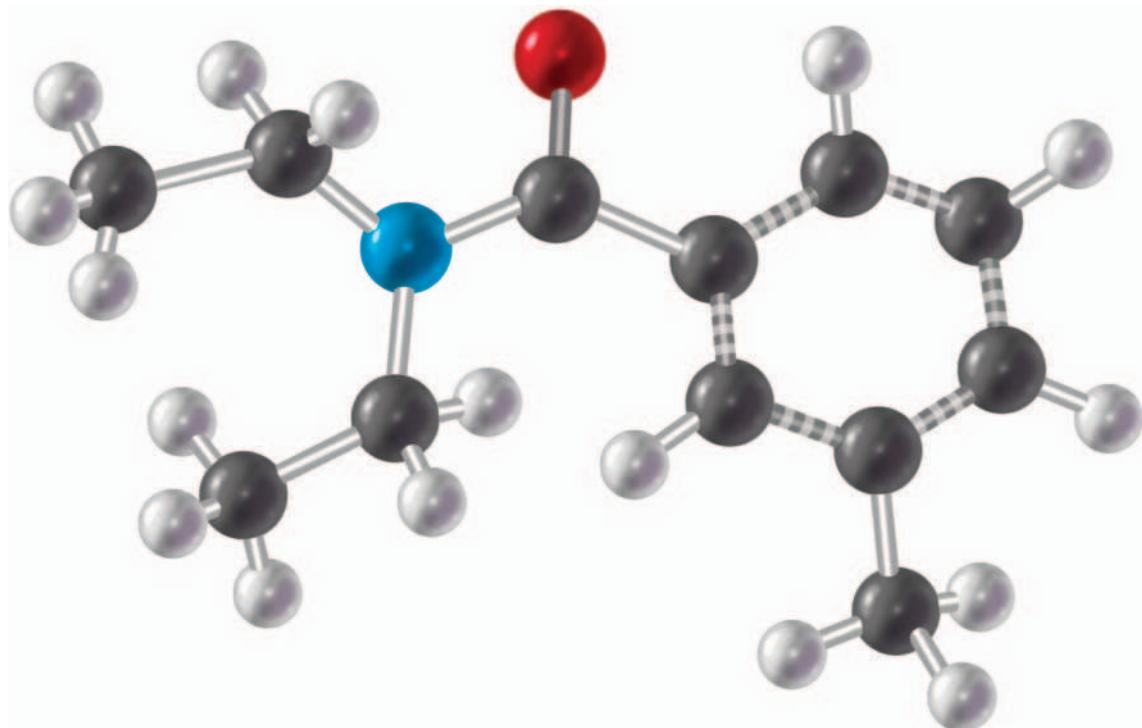
**K
E
Y
F
A
C
T
S**

OVERVIEW

N,N-diethyl-3-methylbenzamide (en-en-dye-ETH-el-three-METH-el-ben-ZA-mid) is the most commonly used insect repellant in the world. It is probably better known by its common name of DEET. DEET is applied to the skin and clothing to ward off biting insects such as mosquitoes, ticks, chiggers, and fleas. Although it does not kill insects, DEET repels them for several hours after being applied.

N,N-diethyl-3-methylbenzamide is a colorless amber liquid with a faint odor. The compound can exist in any one of three isomers in which the two groups attached to the benzene ring are next to each other (ortho), separated by one carbon (meta), or opposite each other in the ring (para). Although all three isomers are effective as insect repellants, the meta isomer is more effective than the ortho and para isomers and constitutes about 95 percent of the commercial product.

DEET was discovered by researchers at the U.S. Department of Agriculture and patented by the U.S. Army in 1946



N,N diethyl 3 methyl benzamide. Red atom is oxygen; white atoms are hydrogen; black atoms are carbon; and blue atom is nitrogen. Gray stick indicates a double bond; striped sticks show a benzene ring.

PUBLISHERS RESOURCE GROUP

for use on military personnel working in insect-infested areas. It was not made available to the general public until 1957. Experts now regard DEET as the “gold standard” of insect repellants. It is available in a variety of formulations, including solutions, lotions, creams, gels, aerosol and pump sprays, and impregnated towelettes under names such as Hour Guard 8 and Hour Guard 12, DEET Plus, Sawyers Gold, Off (in many varieties), Ben’s Backyard and Ben’s Wilderness, Cutter (in many varieties), and Repel (in many varieties).

HOW IT IS MADE

DEET is made commercially by treating *m*-toluic acid ($C_6H_4CH_3COOH$) with thionyl chloride ($SOCl_2$). The product of this reaction, *m*-toluoyl chloride ($C_6H_4CH_3COOC$) is then treated with diethylamine ($C_2H_5)_2NH$) to obtain N,N-diethyl-3-methylbenzamide.

Interesting Facts

- The U.S. Environmental Protection Agency (EPA) has estimated that more than two hundred million people around the world use products containing N,N-diethyl-3-methylbenzamide.

COMMON USES AND POTENTIAL HAZARDS

The mechanism by which DEET repels insects is not completely understood. Current theories suggest that the compound blocks receptors on a mosquito's antennae that help it to locate carbon dioxide and lactic acid, given off in skin perspiration and the breath. Lacking these chemical clues, insects are unable to locate their prey (humans and other animals).

By warding off biting insects, DEET protects against the diseases they carry. Mosquitoes, for example, carry diseases such as malaria, one of the most serious diseases in the world, responsible for an estimated three millions deaths a year; encephalitis, an infection that causes inflammation and swelling of the brain; and West Nile virus, an organism that affects the central nervous system and poses a serious threat to both humans and other animals. Ticks carry Lyme disease, an infection spread by the deer tick that causes a skin rash, joint pain, and flu-like symptoms that can develop into a debilitating and permanent health problem if not treated early.

DEET is a very safe product when used as directed. It is absorbed rapidly through the skin, with up to 56 percent of the compound penetrating the skin in a six-hour period. Within twelve hours of application, DEET is metabolized by the liver and excreted in the urine. If inhaled or swallowed, however, DEET poses serious health hazards. It may cause headaches, seizures, and swelling of the brain, although these symptoms are very rare among users of the product. Products containing DEET are not recommended on babies under the age of two months, and are now required by the EPA

Words to Know

ISOMER One of two or more forms of a chemical compound with the same molecular formula, but different structural

formulas and different chemical and physical properties.

to carry a warning to avoid spraying the product directly into the eyes or open wounds.

FOR FURTHER INFORMATION

"Basic Facts About DEET and DEET Based Insect Repellents." Consumer Specialty Products Association.
http://www.deet.com/deet_fact_sheet.htm (accessed on October 7, 2005).

"DEET."

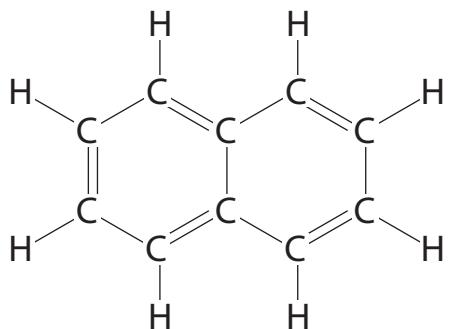
<http://www.chem.ox.ac.uk/mom/InsectRepellents/DEET.htm> (accessed on October 7, 2005).

Fradin, Mark S., M.D. "Mosquitoes and Mosquito Repellants." *Annals of Internal Medicine*. June 1, 1998: 931 940. Also available online at
<http://www.annals.org/cgi/content/full/128/11/931> (accessed on October 7, 2005).

Qiu H., H. W. Jun, and J. W. McCall. "Pharmacokinetics, formulation, and safety of insect repellent N,N diethyl 3 methylbenzamide (DEET): a review." *Journal of the American Mosquito Control Association*. March 1998: 12 27. Also available online at
http://www.ncbi.nlm.nih.gov/entrez/query.fcgi?cmd=Retrieve&db=PubMed&list_uids=98261685&dopt=Citation (accessed on October 7, 2005).

"Review of the Toxicology Literature for the Topical Insect Repellent Diethyl m toluamide (DEET)." Department of Health Toxicology Unit (United Kingdom).
<http://www.advisorybodies.doh.gov.uk/pdfs/reviewofdeet.pdf> (accessed on October 7, 2005).

See Also Carbon Dioxide; Lactic Acid

**OTHER NAMES:**

None

FORMULA:

C₁₀H₈

ELEMENTS:

Carbon, hydrogen

COMPOUND TYPE:

Aromatic hydrocarbon (organic)

STATE:

Solid

MOLECULAR WEIGHT:

128.17 g/mol

MELTING POINT:

80.26°C (176.5°F)

BOILING POINT:

217.9°C (424.2°F)

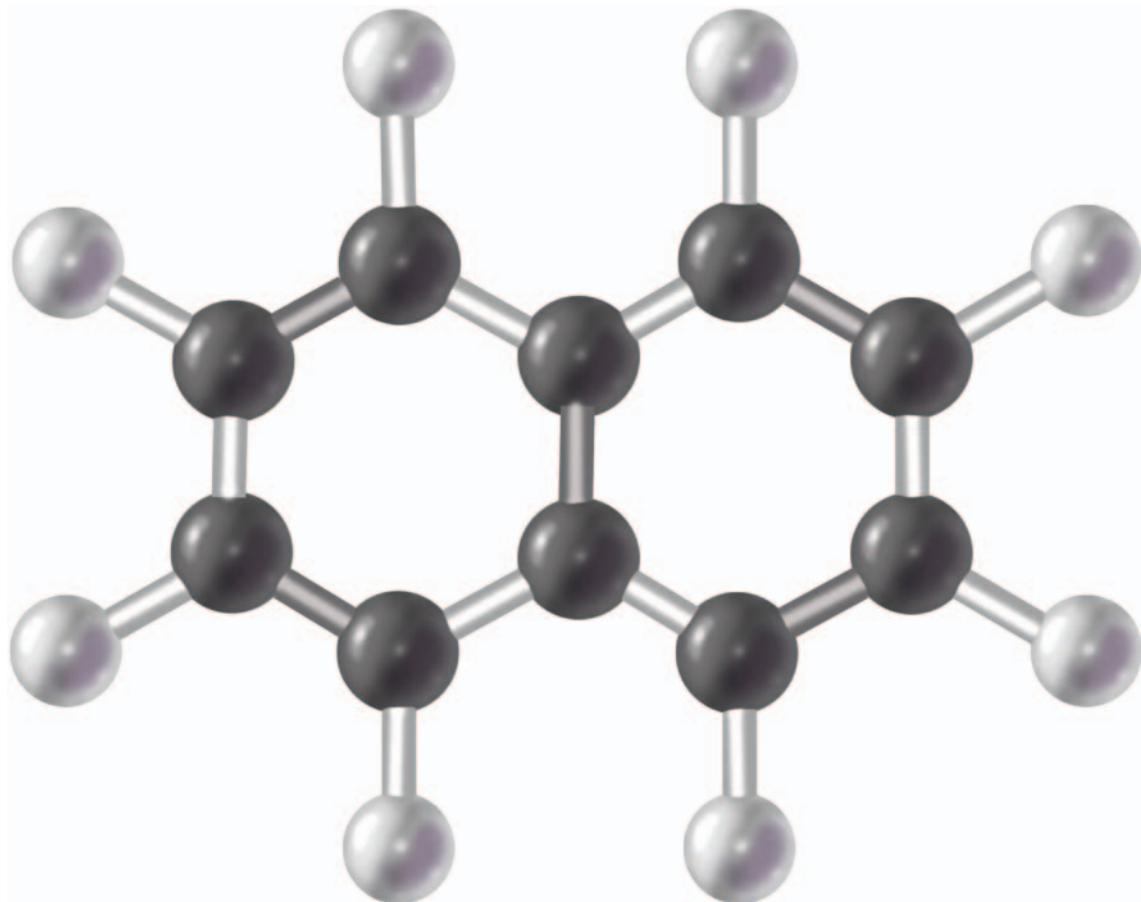
SOLUBILITY:

Insoluble in water;
soluble in ethyl
alcohol; very soluble
in ether and acetone

K E Y F A C T S**OVERVIEW**

Naphthalene (NAF-thuh-leen) is a white crystalline volatile solid with a characteristic odor often associated with mothballs. The compound sublimes (turns from a solid to a gas) slowly at room temperature, producing a vapor that is highly combustible. Naphthalene was first extracted from coal tar in 1819 by English chemist and physician John Kidd (1775–1851). Coal tar is a brown to black thick liquid formed when soft coal is burned in an insufficient amount of air. It consists of a complex mixture of hydrocarbons, similar to that found in petroleum. Kidd's extraction of naphthalene was of considerable historic significance because it demonstrated that coal had other important applications than its use as a fuel. It could also be utilized as the source of chemical compounds with a host of important commercial and industrial uses. Naphthalene's chemical structure was determined by the German chemist Richard August Carl Emil Erlenmeyer (1825–1909). Erlenmeyer showed that the naphthalene molecule consists of two benzene molecules joined to each other.

NAPHTHALENE



Naphthalene. White atoms are hydrogen and black atoms are carbon. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

HOW IT IS MADE

Naphthalene occurs naturally in petroleum and coal tar. It is extracted from either by heating the raw material to a temperature of 200°C to 250°C (392°F to 482°F), producing a mixture of hydrocarbons known as middle oil. The middle oil is then distilled to separate its individual components from each other, one of which is naphthalene. The naphthalene produced in this process is purified by washing it in a strong acid and then in a sodium hydroxide solution and purified by steam distillation.

COMMON USES AND POTENTIAL HAZARDS

In 2000, about 100,000 metric tons (110,000 short tons) of naphthalene was produced in the United States. About 60

Interesting Facts

- The name “naphthalene” comes from the Persian word *naphtha*, meaning “oil” or “pitch.”
- One use of naphthalene is as an insecticide. Nonetheless, one type of insect, the Formosan termite, uses naphthalene to build its nests. Scientists have not yet determined the source from which the termites get naphthalene or why they are immune to its otherwise toxic properties.
- Naphthalene becomes an environmental contaminant when it leaches into the soil. In 2003, however, scientists discovered bacteria capable of decomposing naphthalene in the soil, reducing the environmental hazards it poses.
- One of the most terrible weapons used during the Vietnam War of the 1960s and 1970s was a mixture of naphthalene and palmitate called napalm that ignited and burned at a very high temperature causing widespread destruction of plant life, structures, and human life.

percent of that amount was used for the synthesis of phthalic anhydride [$C_6H_4(CO)_2O$], a compound used as the starting point in the manufacture of a host of products, including a variety of dyes, resins, lubricants, insecticides, and a number of other products. At one time, naphthalene was used widely as a moth repellent in mothballs and moth flakes and in the manufacture of insecticides and fungicides. Demand for these applications has decreased, however, with the introduction of a class of compounds known as the chlorinated hydrocarbons. For example, one of those compounds, *p*-dichlorobenzene, is now the primary ingredient, rather than naphthalene, in mothballs and moth flakes.

Some other uses of naphthalene include:

- As an ingredient in the manufacture of the explosive known as smokeless gunpowder;
- In the manufacture of scintillation counters, devices used to detect and measure radioactivity;

Words to Know

DISTILLATION A process of separating two or more substances by boiling the mixture of which they are composed and

condensing the vapors produced at different temperatures.

- In the production of lubricants;
- In the manufacture of various types of dyes and coloring agents;
- As an antiseptic and antihelminthic (a substance used to kill disease-causing worms);
- As a deodorant in toilets, diaper pails, and other settings; and
- As a preservative.

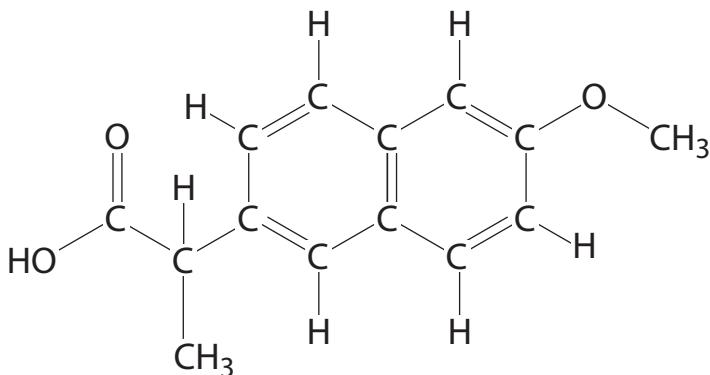
Naphthalene can be toxic by ingestion, inhalation, or absorption through the skin when humans are exposed to it in more than a passing way. In small amounts it may produce symptoms such as nausea, vomiting, headache, and fever. In larger amounts, it may induce more serious problems, including liver damage, blindness, coma, convulsions, and death. One medical problem that has been associated with exposure to large amounts of naphthalene is hemolytic anemia, a condition in which red blood cells are destroyed, resulting in fatigue, restlessness, jaundice, loss of appetite, and, in more serious cases, kidney failure.

FOR FURTHER INFORMATION

"Background and Environmental Exposures to Naphthalene, 1 Methylnaphthalene, and 2 Methylnaphthalene in the United States." Agency for Toxic Substances and Disease Registry. http://www.atsdr.cdc.gov/toxprofiles/tp67_c2.pdf (accessed on October 17, 2005).

Milius, Susan. "Termites Use Mothballs in Their Nests." *Science News* (May 2, 1998): 228.

"Naphthalene Chemical Backgrounder." National Safety Council. <http://www.nsc.org/ehc/chemical/Naphthal.htm> (accessed on October 17, 2005).



OTHER NAMES:
 (S)-6-methoxy-
 α -methyl-
 2-naphthaleneacetic
 acid; *d*-2-(6-methoxy-
 2-naphthyl)-propionic
 acid

FORMULA:
 $C_{14}H_{14}O_3$

ELEMENTS:
 Carbon, hydrogen,
 oxygen

COMPOUND TYPE:
 Carboxylic acid
 (organic)

STATE:
 Solid

MOLECULAR WEIGHT:
 230.26 g/mol

MELTING POINT:
 155°C (311°F)

BOILING POINT:
 Not applicable

SOLUBILITY:
 Insoluble in water;
 soluble in methyl
 alcohol and
 chloroform

KEY FACTS

OVERVIEW

Naproxen (nah-PROK-sin) is a white to off-white odorless crystalline solid sold under a variety of commercial names, including Aleve®, Anaprox®, Bonyl®, Calosen®, Diocodal®, Naprosyn®, Naprelan®, Proxen®, and Veradol®. It is a nonsteroidal anti-inflammatory drug (NSAID) used as an analgesic (pain reliever) and antipyretic (fever-reducing compound). The compound is also available as a sodium salt, which is more readily absorbed in the gastrointestinal tract. Nonsteroidal anti-inflammatory drugs are compounds used to reduce pain, fever, and inflammation without the use of steroids (thus, nonsteroidal) by inhibiting the action of an enzyme needed to produce these results in the body. Other examples of NSAIDs are aspirin, ibuprofen, and prednisone.

HOW IT IS MADE

Naproxen is prepared by treating 2-methoxynaphthalene with a derivative of propanoic acid (CH_3CH_2COOH).

NAPROXEN



Naproxen. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. Striped sticks show benzene rings. PUBLISHERS RESOURCE GROUP

2-methoxynaphthalene has a structure very similar to that of naproxen and requires only the addition of the propanoyl group (-CH₂CH₂COOH) provided by propanoic acid. The reaction is fairly straightforward but includes one important consideration. Naproxen is a chiral compound,

Interesting Facts

- The U.S. Food and Drug Administration first approved the sale of naproxen as an over-the-counter medication (one that does not require a prescription) in 1994.

meaning that it can exist in one of two isomeric forms. The two forms are a “right-handed” form (designated as the R form) and a “left-handed” form (designated as the S form). As is the case with many organic compounds, the two chiral forms of naproxen have very different biological activities. Specifically, the S-form of naproxen is 28 times as effective as the R-form in achieving the analgesic, antipyretic, and anti-inflammatory results expected from the compound. Early methods for preparing naproxen resulted in the formation of a racemic mixture of R and S forms. A racemic mixture is one that contains both of the isomeric forms in which a compound can occur. The two parts of the racemic mixture, the R and S forms, then had to be separated from each other. Researchers have now developed a method of manufacture that produces essentially pure S naproxen, avoiding the time-consuming and costly process of separation previously required.

COMMON USES AND POTENTIAL HAZARDS

Naproxen is commonly used to treat the pain and stiffness caused by conditions such as arthritis, inflammation of the joints, gout, tendonitis, and bursitis. It achieves this result by inhibiting the activity of an enzyme called cyclooxygenase 2 (COX-2). COX-2 facilitates the production of a class of compounds in the body known as prostaglandins. Prostaglandins are produced when the body is injured, causing the pain and inflammation usually associated with an injury. Naproxen blocks the active site on a COX-2 molecule that catalyzes the formation of COX-2 molecules with the result: no COX-2; no prostaglandins; no pain and inflammation. Interestingly enough, a very similar enzyme called cyclooxygenase

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

ISOMER One of two or more forms of a chemical compound with the same

molecular formula, but different structural formulas and different chemical and physical properties.

COX-2 also exists in the body. COX-1 , however, has a different set of functions that have nothing to do with the production of pain and inflammation. Since naproxen inhibits COX-2 , but not COX-1 , it is known as a selective inhibitor NSAID.

A number of relatively minor side effects have been reported for naproxen, including headache, dizziness, drowsiness, itching, ringing in the ears, and minor gastrointestinal discomfort that may include heartburn, abdominal pain, nausea, constipation, and diarrhea. More serious problems have also been reported, including gastrointestinal complications such as bleeding and ulceration (bleeding sores) of the stomach and intestines. In some cases, the drug has been responsible for the hospitalization of users.

Care needs to be taken when combining naproxen with other medications. Known adverse drug interactions can occur with aspirin, methotrexate, ACE inhibitors (for high blood pressure), furosemide, lithium, and warfarin (a blood thinner). An overdose of naproxen may cause dizziness, drowsiness, and gastrointestinal problems. High blood pressure, kidney failure, and coma may occur, but are rare.

FOR FURTHER INFORMATION

Buschmann, Helmut, et al. *Analgesics: From Chemistry and Pharmacology to Clinical Application*. New York: Wiley VCH, 2002.

“Naproxen.” World of Molecules.

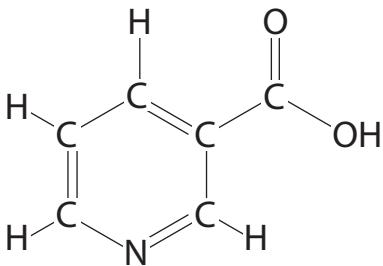
<http://www.worldofmolecules.com/drugs/naproxen.htm>
(accessed on October 17, 2005).

Omudhome, Ogbu. "Naproxen Center."

<http://www.medicinenet.com/naproxen/article.htm> (accessed on October 17, 2005).

"Product Monograph: Anaprox®; Anaprox® DS." Roche.

<http://www.rochecanada.com/pdf/Anaprox%20PM%20E.pdf>
(accessed on October 17, 2005).



OTHER NAMES:	Nicotinic acid; 3-pyridinecarboxylic acid; vitamin B ₃
FORMULA:	C ₆ H ₅ NO ₂
ELEMENTS:	Carbon, hydrogen, nitrogen, oxygen
COMPOUND TYPE:	Carboxylic acid (organic)
STATE:	Solid
MOLECULAR WEIGHT:	123.11 g/mol
MELTING POINT:	236.6°C (457.9°F)
BOILING POINT:	Not applicable; sublimes above its melting point
SOLUBILITY:	Slightly soluble in water, ethyl alcohol, and ether

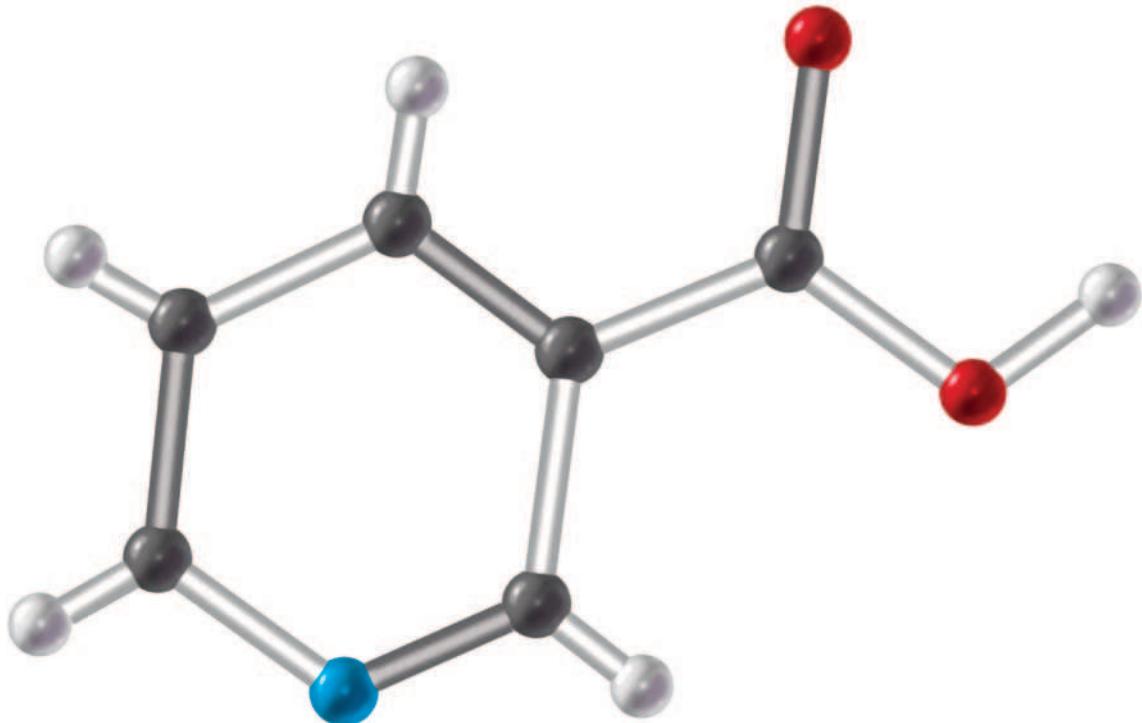
KEY FACTS

Niacin

OVERVIEW

Niacin (NYE-uh-sin) is a B vitamin (vitamin B₃) that is essential to cell metabolism. It occurs in two forms, nicotinic acid and niacinamide, also called niacinamide. The only structural difference between the two compounds is that a hydroxyl group (-OH) in nicotinic acid is replaced by an amino group (-NH₂) group in niacinamide. Lack of niacin causes a disease called pellagra. Pellagra was common throughout human history among poor people whose diet consisted almost entirely of corn products. Those corn products did not supply adequate amounts of niacin, causing symptoms such as diarrhea, scaly skin sores, inflamed mucous membranes, weakness, irritability, and mental delusions. In some cases, people with niacin deficiency develop reddish sores and rashes on their faces. Mental hospitals were full of people who seemed crazy, but who were actually suffering from a dietary deficiency. Thousands of people died from pellagra every year.

Nicotinic acid was first isolated by the Polish-American biochemist Casimir Funk (1884-1967) in 1912. At the time,



Niacin. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; blue atom is nitrogen. Gray sticks indicate double bonds.

PUBLISHERS RESOURCE GROUP

Funk was attempting to find a cure for another dietary disease known as beriberi. Since nicotinic acid had no effect on beriberi, he abandoned his work with that compound. It was left, then, to the Austrian-American physician Joseph Goldberger (1874-1929) to find the connection between nicotinic acid and deficiency diseases. In 1915, Goldberger conducted a series of experiments with prisoners in a Mississippi jail and found that he could produce pellagra by altering their diets. He concluded that the disease was caused by the absence of some factor, which he called the P-P (for pellagra-preventative) factor. The chemical structure of that factor was then discovered in 1937 by the American biochemist Conrad Arnold Elvehjem (1901-1962), who cured the disease in dogs by treating them with nicotinic acid.

HOW IT IS MADE

Niacin is synthesized naturally in the human body beginning with the amino acid tryptophan. Tryptophan occurs naturally in a number of foods, including dairy products,

Interesting Facts

- Early investigators chose to use the name niacin rather than nicotinic acid because the later term sounds too much like “nicotine.” They did not want to give the impression that a person can get the niacin they need by smoking tobacco, which contains nicotine.
- Some good sources of niacin include organ meats, such as kidney, liver, and heart; chicken; fish, such as salmon and tuna; milk and dairy products; brewer’s yeast; dried beans and legumes; nuts and seeds; broccoli and asparagus; whole grains; dates; mushrooms; tomatoes; sweet potatoes; and avocados.
- Corn is also a good source of niacin. However, people with corn-based diets are at risk for developing pellagra. The reason for this contradiction is that the niacin in corn is chemically inactivated. Corn must be treated with an alkaline material to convert the niacin it contains to a free form that the body can use. Native Americans who treated their corn products with ley, limestone, wood ashes, or other alkaline materials did not suffer from pellagra.

beef, poultry, barley, brown rice, fish, soybeans, and peanuts. People whose diet consists mainly of corn products do not ingest adequate amounts of tryptophan, so their bodies are unable to make the niacin they need to avoid developing pellagra. It takes about 60 milligrams of tryptophan to produce 1 mg of niacin.

COMMON USES AND POTENTIAL HAZARDS

Niacin plays a number of essential roles in the body. It is necessary for cell respiration; metabolism of proteins, fats, and carbohydrates; the release of energy from foods; the secretion of digestive enzymes; the synthesis of sex hormones; and the proper functioning of the nervous system. It is also involved in the production of serotonin, an essential

Words to Know

METABOLISM Process that includes all of the chemical reactions that occur in cells by which fats, carbohydrates, and other compounds are broken down to produce energy and the compounds needed to build new cells and tissues.

MUCOUS MEMBRANES Tissues that line the moist inner lining of the digestive,

respiratory, urinary, and reproductive systems.

NEUROTRANSMITTER A chemical that carries nerve transmissions from one nerve cell to an adjacent nerve cell.

SUBLIME Change in state from a solid directly to a gas.

neurotransmitter in the brain. Niacin deficiency disorders occur as the result of an inadequate diet, consuming too much alcohol, and among people with certain types of cancer and kidney diseases. Physicians treat niacin deficiency diseases by prescribing supplements of 300 to 1,000 milligrams per day of the vitamin. Overdoses of niacin can cause a variety of symptoms, including itching, burning, flushing, and tingling of the skin.

FOR FURTHER INFORMATION

Eades, Mary Dan. *The Doctor's Complete Guide to Vitamins and Minerals*. New York: Dell, 2000.

“Niacin (Nicotinic Acid).” PDRHealth.

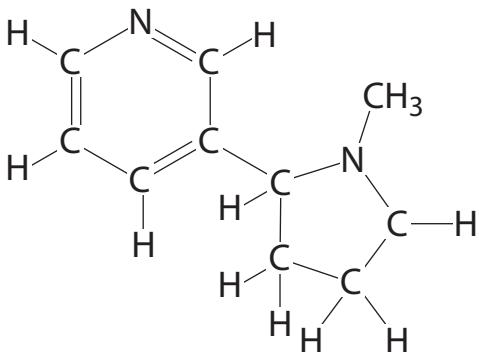
http://www.pdrhealth.com/drug_info/nmdrugprofiles/nutsupdrugs/nia_0184.shtml (accessed on October 20, 2005).

“Niacin Deficiency.” *The Merck Manual*.

<http://www.merck.com/mrkshared/mmanual/section1/chapter3/31.jsp> (accessed on October 20, 2005).

“Nicotinic Acid.” Gondar Design Science.

<http://www.purchon.com/biology/nicotinic.htm> (accessed on October 20, 2005).

**OTHER NAMES:**

r(S)-3-(1-methyl-2-pyrrolidinyl)pyridine;
1-methyl-2-(3-pyridyl)pyrrolidine

FORMULA:
 $C_8H_{10}N_2$

ELEMENTS:
Carbon, hydrogen,
nitrogen

COMPOUND TYPE:
Alkaloid (organic)

STATE:
Liquid

MOLECULAR WEIGHT:
162.23 g/mol

MELTING POINT:
−79°C (−110°F)

BOILING POINT:
247°C (477°F)

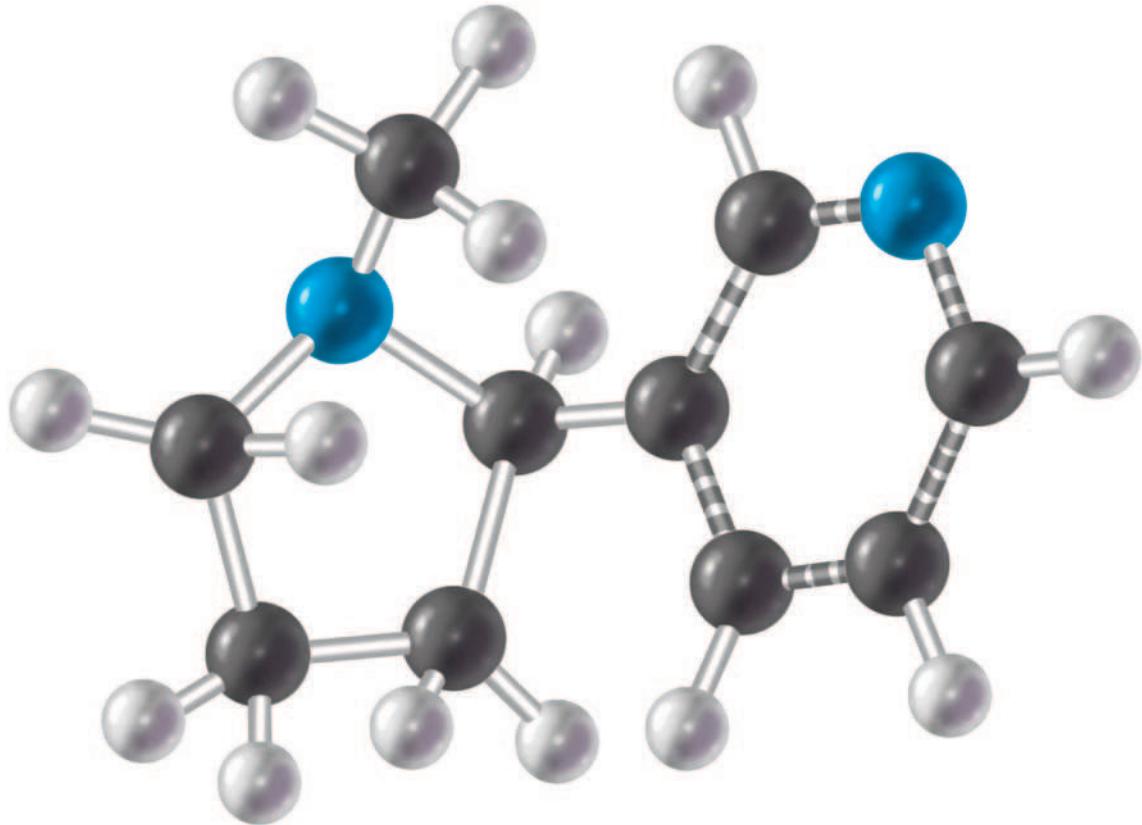
SOLUBILITY:
Miscible with water;
very soluble in ethyl
alcohol, ether, and
chloroform

K E Y F A C T S**OVERVIEW**

Nicotine (NIK-uh-teen) is a thick, colorless to yellow, oily liquid with a bitter taste that turns brown when exposed to air. It occurs in high concentrations in the leaves of tobacco plants and in lower concentrations in tomatoes, potatoes, eggplants, and green peppers. Nicotine gets its name from the tobacco plant, *Nicotiana tabacum*, which, in turn, was named in honor of the French diplomat and scholar Jean Nicot (1530–1600), who introduced the use of tobacco to Paris. Nicotine's correct chemical structure was determined in 1843 by the Belgian chemist and physicist Louise Melsens (1814–1886) and the compound was first synthesized by the research team of A. Pictet and A. Rotschy in 1904.

HOW IT IS MADE

Nicotine is extracted by soaking the stems and leaves of the tobacco plant in water for about twelve hours. After that period of time, the nicotine in the tobacco has dissolved in



Nicotine. White atoms are hydrogen; black atoms are carbon; and blue atoms are nitrogen. PUBLISHERS RESOURCE GROUP

the water and can be extracted in a variety of ways. In one process, the water solution of nicotine is mixed with ether or chloroform, in which the nicotine is more soluble. The nicotine moves from the water layer to the ether or chloroform layer, from which it can be removed by evaporation.

COMMON USES AND POTENTIAL HAZARDS

The best-known application of nicotine is in tobacco products used for smoking and chewing. The actual nicotine content of tobacco products varies considerably, but, on average, ranges from about 15 to 25 milligrams per cigarette. Nicotine is also available in a number of products designed to help people stop smoking, such as nicotine gums and nicotine patches.

Nicotine was often used by farmers and gardeners as an insecticide and a fumigant in the past. Perhaps the best

Interesting Facts

- It takes only about seven seconds after nicotine is ingested before the chemical reaches the human brain.
- Nicotine is one of the most widely used addictive drugs in the United States.

known of these in the United States was an insecticide known as Black Leaf 40, a 40 percent solution of nicotine sulfate in water. The use of Black Leaf 40 and most other nicotine-containing insecticides has, to a large extent, been discontinued because of the toxic nature of the compound. The risk it posed to human users was greater than its value as an agricultural product.

Nicotine is a highly addictive substance. For that reason, people have difficulty stopping smoking or chewing tobacco products even when they recognize the health hazards posed by the compound. Smokers depend on nicotine to give them a burst of energy, since it stimulates the heart rate and quickens blood flow. Once a person becomes addicted to the use of nicotine, it requires larger doses of the compound to produce comparable effects. A 1998 U.S. government report issued by then-Surgeon General C. Everett Koop found that the addictive properties of nicotine are comparable to those of heroin and cocaine.

When a person uses tobacco, nicotine is quickly absorbed through respiratory tissues, the skin, and the gastrointestinal tract. The actual amount of nicotine absorbed by the body depends on a number of factors, including the type of tobacco being smoked and the presence or absence of a filter on the cigarette. After entering the body, nicotine flows through the bloodstream and across the blood brain barrier. Levels of the stimulating hormone adrenaline increase, as do blood sugar levels, respiration rates, blood pressure, and heart rate. Nicotine can make small arteries constrict, putting strain on the heart and raising blood pressure. If a person already has clogged arteries, this effect may cause heart pain (angina) or a heart attack.

Words to Know

MISCIBLE Able to be mixed; especially applies to the mixing of one liquid with another.

NEUROTRANSMITTER A chemical that carries nerve transmissions from one nerve cell to an adjacent nerve cell.

Although nicotine is a stimulant, it may induce muscle relaxation, depending on the user's physical state. It has also been shown to decrease one's appetite, speed up metabolism, and increase levels of dopamine, a mood-altering chemical in the brain that induces feelings of pleasure. Low levels of dopamine play a role in the development of Parkinson's disease. Research has shown that smokers, with higher levels of dopamine, have a reduced risk of the disease. Women who are pregnant are advised not to use any product containing nicotine. Nicotine in any form is harmful to an unborn child. It rapidly crosses the placenta and enters the fetus's body.

Nicotine is a highly toxic poison, which explains its former popularity as a pesticide. In high doses, it can be lethal. Low doses of nicotine can cause dizziness, nausea, and vomiting. Symptoms of acute nicotine poisoning may include a burning sensation in the mouth, more severe nausea and vomiting, diarrhea, heart palpitations, fluid in the lungs, seizures, coma, and death. People who smoke while receiving nicotine replacement therapy are at risk of nicotine poisoning.

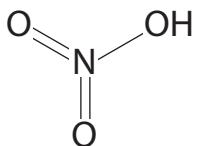
FOR FURTHER INFORMATION

"Acute Nicotine Poisoning." *Mosby's Medical, Nursing, and Allied Health Dictionary*. 5th edition. St. Louis: Mosby, 1998.

"Facts about Nicotine and Tobacco Products." National Institute on Drug Abuse.
http://www.drugabuse.gov/NIDA_Notes/NNVol13N3/tearoff.html (accessed on October 20, 2005).

"Nicotine." International Labour Organization.
http://www.ilo.org/public/english/protection/safework/cis/products/icsc/dtasht/_icsc05/icsco519.htm (accessed on October 20, 2005).

"Nicotine (Black Leaf 40) Chemical Profile 4/85." Pesticide Management Education Program, Cornell University.
[http://pmep.cce.cornell.edu/profiles/insect mite/
mevinphos propargite/nicotine/insect prof nicotine.html](http://pmep.cce.cornell.edu/profiles/insect%20mite/mevinphos%20propargite/nicotine/insect%20prof%20nicotine.html)
(accessed on October 20, 2005).

**OTHER NAMES:**

Aqua fortis;
engraver's acid;
azotic acid

FORMULA:

HNO_3

ELEMENTS:

Hydrogen, nitrogen,
oxygen

COMPOUND TYPE:

Inorganic acid

STATE:

Liquid

MOLECULAR WEIGHT:

63.01 g/mol

MELTING POINT:

-41.6°C (-42.9°F)

BOILING POINT:

83°C (180°F);
decomposes

SOLUBILITY:

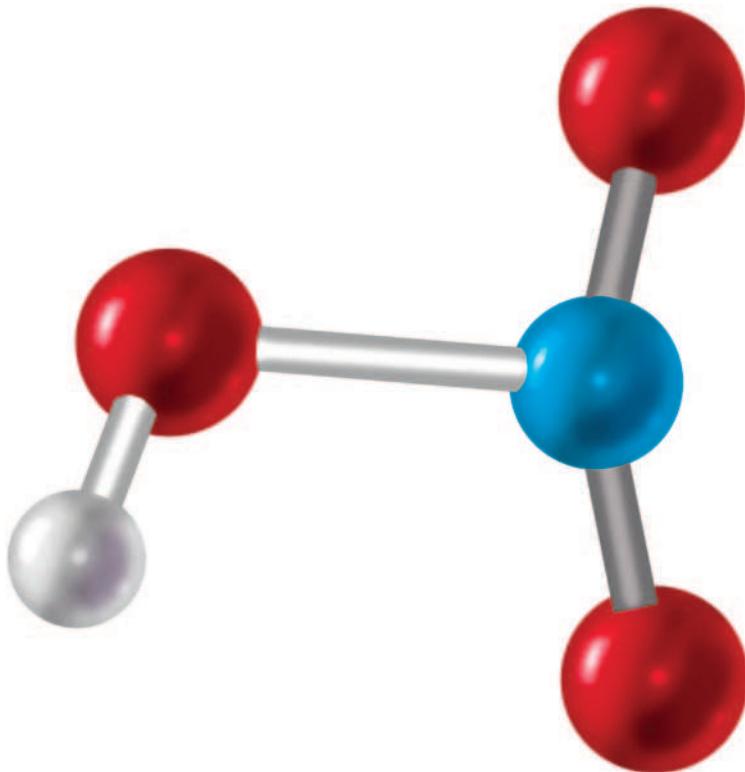
Miscible with water;
decomposes in ethyl
alcohol; reacts
violently with most
organic solvents

KEY FACTS**OVERVIEW**

Nitric acid (NYE-trik AS-id) is a colorless to yellowish liquid with a distinctive acrid (biting), suffocating, or choking odor. The acid tends to decompose when exposed to light, producing nitrogen dioxide (NO_2), itself a brownish gas. The yellowish tinge often observed in nitric acid is caused by the presence of small amounts of the nitrogen dioxide. Nitric acid is one of the strongest oxidizing agents known and attacks almost all metals with the notable exceptions of gold and platinum.

Nitric acid has been known to scholars for many centuries. Probably the earliest description of its synthesis occurs in the writings of the Arabic alchemist Abu Musa Jabir ibn Hayyan (c. 721-c. 815), better known by his Latinized name of Geber. The compound was widely used by the alchemists, although they knew nothing of its chemical composition. It was not until the middle of the seventeenth century that an improved method for making nitric acid was invented by German chemist Johann Rudolf Glauber (1604-1670). Glauber produced the acid by adding concentrated sulfuric acid (H_2SO_4)

Nitric acid. Red atoms are oxygen; white atom is hydrogen; and blue atom is nitrogen. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP



to saltpeter (potassium nitrate; KNO_3). A similar method is still used for the laboratory preparation of nitric acid, although it has little or no commercial or industrial value.

The chemical nature and composition of nitric acid were first determined in 1784 by the English chemist and physicist Henry Cavendish (1731–1810). Cavendish applied an electric spark to moist air and found that a new compound—nitric acid—was formed. Cavendish was later able to determine the acid's chemical and physical properties and its chemical composition. The method of preparation most commonly used for nitric acid today was one invented in 1901 by the Russian-born German chemist Friedrich Wilhelm Ostwald (1853–1932). The Ostwald process involves the oxidation of ammonia over a catalyst of platinum or a platinum-rhodium mixture.

Today, nitric acid is one of the most important chemical compounds used in industry. It ranks number thirteen among all chemicals produced in the United States each year. In 2005, about 6.7 million metric tons (7.4 million short tons) of the compound were produced in the United States.

Interesting Facts

- Alchemists called nitric acid *aqua fortis*, a term that means “strong water.”
 - Nitric acid is a component of acid rain, a form of pol-
- lution that results when substances such as nitrogen oxides react with water, oxygen, and other chemicals in the atmosphere.

HOW IT IS MADE

Although several methods for the preparation of nitric acid are theoretically available, only one finds much commercial use: the direct oxidation of ammonia, an updated and improved version of the traditional Ostwald process. In this method, ammonia is heated and reacted with air over a catalyst, most commonly a mixture of rhodium and platinum metals. That reaction results in the formation of nitric oxide (NO), which is then converted to nitrogen dioxide (NO₂). The nitrogen dioxide reacts with water to form nitric acid.

COMMON USES AND POTENTIAL HAZARDS

The most common use for nitric acid is in the manufacture of ammonium nitrate, which, in turn, is used primarily as a fertilizer. About three-quarters of all nitric acid produced in the United States is used in fertilizers. The second most important application, accounting for about 10 percent of all nitric acid produced, is in the production of adipic acid [COOH(CH₂)₃COOH], used in the manufacture of nylon, polyurethanes, and other synthetic plastics. Nitric acid is also used to make a variety of metal nitrates and for the cleaning of metals. Small amounts of the compound are used for a variety of other applications, including:

- In the manufacture of explosives and fireworks;
- As a laboratory reagent in commercial, industrial, and academic research laboratories;
- In the processing of nuclear fuels;
- For the etching of metals; and
- In the manufacture of certain types of dyes.

Words to Know

ALCHEMY An ancient field of study from which the modern science of chemistry evolved.

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

MISCIBLE Able to be mixed; especially applies to the mixing of one liquid with another.

OXIDATION A chemical reaction in which oxygen reacts with some other substance or, alternatively, in which some substance loses electrons to another substance, the oxidizing agent.

Nitric acid is a highly toxic material. It attacks and destroys skin and other tissues, leaving a distinctive yellow scar caused by the destruction of proteins in the skin or tissue. If swallowed, inhaled, or spilled on the skin, it can cause a number of effects, including severe corrosive burns to the mouth, throat, and stomach; severe irritation or burning of the upper respiratory system, including nose, mouth, and throat; damage to the lungs; severe breathing problems; and burns to the eye surface, conjunctivitis, and blindness. In the most severe cases, the acid can cause death.

FOR FURTHER INFORMATION

“Concentrated Nitric Acid (70%).” International Chemical Safety Cards.

<http://www.cdc.gov/niosh/ipcsneng/nengo183.html> (accessed on October 20, 2005).

“Material Safety Data Sheet.” Hill Brothers Chemical Company.

http://hillbrothers.com/msds/pdf/nitric_acid.pdf (accessed on October 20, 2005).

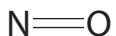
“Nitric Acid.” Greener Industry.

http://www.uyseg.org/greener_industry/pages/nitric_acid/1nitricAcidAP.htm (accessed on October 20, 2005).

“Nitric Acid.” Scorecard. The Pollution Information Site.

http://www.scorecard.org/chemical_profiles/summary.tcl?edf_substance_id=7697%2d37%2d2 (accessed on October 20, 2005).

See Also Ammonia; Ammonium Nitrate; Nitrogen Dioxide; Nitroglycerin



OTHER NAMES:	Nitrogen monoxide
FORMULA:	NO
ELEMENTS:	Nitrogen, oxygen
COMPOUND TYPE:	Nonmetallic oxide (inorganic)
STATE:	Gas
MOLECULAR WEIGHT:	36.01 g/mol
MELTING POINT:	-163.6°C (-262.5°F)
BOILING POINT:	-151.74°C (-241.13°F)
SOLUBILITY:	Slightly soluble in water

KEY FACTS

Nitric Oxide

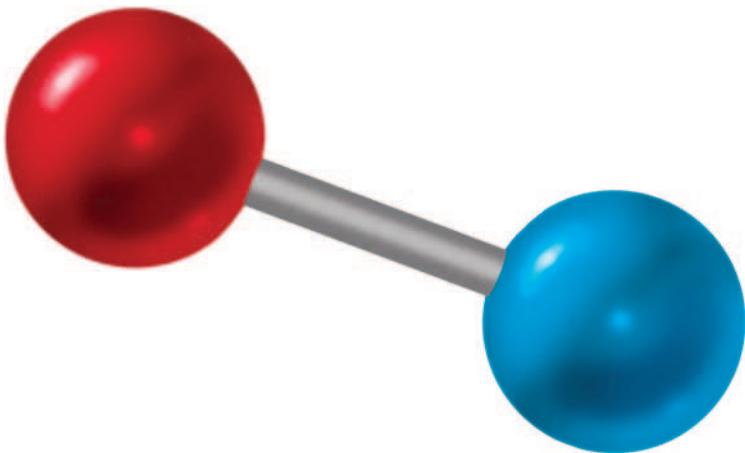
OVERVIEW

Nitric oxide (NYE-trik OK-side) is a sweet-smelling, colorless gas that can be liquefied to make a bluish liquid and frozen to produce a bluish-white snow-like solid. It is one of five oxides of nitrogen, the others being nitrous oxide (N_2O), nitric oxide (NO), dinitrogen trioxide N_2O_3 , and nitrogen dioxide (NO_2). Nitric oxide was first discovered in 1620 by Flemish physician and alchemist Jan Baptista van Helmont (1580–1635 or 1644).

Nitric oxide is used in the production of nitric acid, ammonia, and other nitrogen-containing compounds. It is also formed as a byproduct of the combustion of coal and petroleum products. As such, it is a major contributor to air pollution.

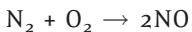
HOW IT IS MADE

Nitrogen and oxygen are the two most abundant gases in the atmosphere. Since both elements are relatively inactive,



Nitrogen monoxide. Red atom is oxygen and blue atom is nitrogen. PUBLISHERS RESOURCE GROUP

they do not combine with each other under normal circumstances. However, the energy provided by lightning strikes causes the reaction of the two elements, producing nitric oxide.



Another source of the energy needed for this reaction is the combustion of coal and oil products used for human activities. For example, the combustion of gasoline in an internal combustion engine produces temperatures in excess of $2,000^{\circ}\text{C}$ (3600°F). These temperatures are sufficient to bring about the reaction between nitrogen and oxygen in a vehicle engine, resulting in the formation of nitric oxide as one product of gasoline combustion. That nitric oxide then passes into the atmosphere and becomes a major component of air pollution.

COMMON USES AND POTENTIAL HAZARDS

The most important industrial use of nitric oxide is in the preparation of other nitrogen-containing compounds, especially nitrogen dioxide (NO_2), nitric acid (HNO_3), and nitrosyl chloride (NOCl). It also finds some application in the bleaching of rayon (a synthetic, or artificially created, fabric) and as a polymerization inhibitor with certain compounds such as propylene and methyl ether. Such compounds have a tendency to react with each other to form large, complex molecules known as polymers.

Nitric oxide is considered an environmental pollutant. It oxidizes readily to form nitrogen dioxide, which, in turn,

Interesting Facts

- In 1992, *Science* magazine named nitric oxide “Molecule of the Year” after scientists discovered that it had several important functions in the body.
- Ferid Murad (1936–), Robert Furchtgott (1916–), and Louis Ignarro (1941–) shared the 1998 Nobel Prize in Physiology or Medicine for their discovery of the role played by nitric oxide in the body’s nervous system.

reacts with moisture in the air to form nitric acid, a component of acid rain. Acid rain is thought to be responsible for a number of environmental problems, including damage to buildings, destruction of trees, and the death of aquatic life. The nitrogen dioxide produced from nitric oxide is also a primary component of photochemical smog, a hazardous haze created by a mixture of pollutants in the presence of sunlight.

Even though it is toxic in the environment, nitric oxide plays several important roles in the human body. Nitric oxide is involved in the process by which messages are transmitted from one nerve cell to the next. It also regulates blood flow by triggering the smooth muscles surrounding blood vessels to relax. This action increases blood flow and lowers blood pressure. Nitric oxide also prevents the formation of blood clots, which can break off and travel to the heart or brain, increasing the risk of heart attack or stroke.

During sexual arousal, nitric oxide increases blood flow to the penis, leading to an erection in a man. The drug Viagra stimulates erections by enhancing the flow of nitric oxide in the penis.

Finally, nitric oxide plays a role in memory and learning. A deficiency of the compound appears to be related to the development of learning problems. On the other hand, an excess of nitric oxide has been implicated in the development of certain diseases, such as Huntington’s chorea, an inherited disorder characterized by unusual body movements

Words to Know

POLYMERIZATION The process of creating a polymer, a compound consisting of very

large molecules made of one or two small repeated units called monomers.

and memory loss, and Alzheimer's disease, a progressive disorder that results in memory loss.

When used to treat a medical condition, nitric oxide is usually administered in the form of a solid or liquid medicine that decomposes in the body, releasing the compound. For example, the drug nitroglycerin is used to treat heart problems. When it enters the bloodstream, nitroglycerin begins to break down, releasing nitric oxide. The nitric oxide causes smooth muscle cells in the heart to relax, relieving the symptoms of angina, chest pain caused by an inadequate flow of blood to the heart. Other types of drugs produce nitric oxide to inhibit the buildup of fatty deposits in blood vessels, which can lead to heart attack and stroke. Patients with pulmonary hypertension, a condition in which the vessels that supply blood to the lungs are constricted, preventing normal oxygen flow, are sometimes given an inhaler with a mixture of nitric oxide and air to open blood vessels to the lungs.

In spite of its many benefits, nitric oxide may also be a health hazard. If inhaled in excessive amounts, it may replace oxygen in the lungs, leading to asphyxia, suffocation resulting from an insufficient supply of oxygen. Research suggests that exposure to low concentrations of the gas over long periods of time may result in lung disease, emphysema, and chronic bronchitis.

FOR FURTHER INFORMATION

Butler, A. R., and R. Nicholson. *Life, Death and Nitric Oxide*. London: Royal Society of Chemistry, 2003.

"Gas Data." Air Liquide.

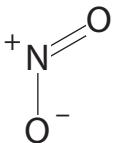
<http://www.airliquide.com/en/business/products/gases/gasdata/index.asp?GasID=44> (accessed on October 20, 2005).

"Nitric Oxide." Reproductive and Cardiovascular Disease Research Group.
<http://www.sgul.ac.uk/depts/immunology/~dash/no/> (accessed on October 20, 2005).

"Nitrogen Oxides." International Programme on Chemical Safety.
<http://www.inchem.org/documents/ehc/ehc/ehc188.htm#SubSectionNumber:2.1.1> (accessed on October 20, 2005).

Stanley, Peter. "Nitric Oxide." Biological Sciences Review (April 2002): 18–20.

See Also Nitric Acid; Nitrogen Dioxide

**OTHER NAMES:**

Dinitrogen tetroxide;
nitrogen peroxide

FORMULA:

NO_2

ELEMENTS:

Nitrogen, oxygen

COMPOUND TYPE:

Nonmetallic oxide
(inorganic)

STATE:

Gas or liquid

MOLECULAR WEIGHT:

46.01 g/mol

MELTING POINT:

-11.2°C (11.8°F)

BOILING POINT:

21.15°C (70.07°F)

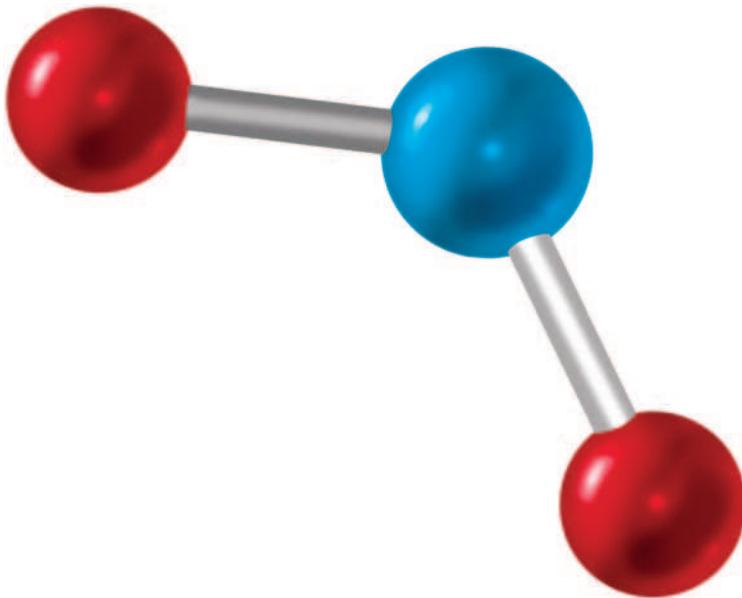
SOLUBILITY:

Reacts with water to
form nitric acid
(HNO_3) and nitrous
acid (HNO_2)

KEY FACTS**OVERVIEW**

Nitrogen dioxide (NYE-truh-jin dye-OK-side) is a toxic reddish-brown gas or yellowish-brown liquid with a pungent, irritating odor. Above 21.15°C (70.07°F), it exists as the reddish-brown gas. Below that temperature, it becomes the yellowish-brown liquid. When liquified under pressure, it forms a fuming brown liquid. The brown liquid is actually a mixture of nitrogen dioxide and dinitrogen tetroxide (N_2O_4), a dimeric form of nitrogen dioxide. A dimer is a molecule that consists of two identical molecules combined with each other. When cooled below -11.2°C (11.8°F), the liquid freezes to form a colorless crystalline solid that consists almost entirely of the dimeric form, N_2O_4 .

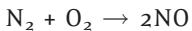
Small amounts of nitrogen dioxide are present naturally in the atmosphere as the result of lightning strikes, volcanic action, forest fires, and bacterial action on dead plants and animals. Much larger amounts are present because of human activities, primarily the combustion of fossil fuels, such as coal and petroleum products.



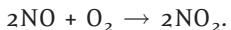
Nitrogen dioxide. Red atoms are oxygen and blue atom is nitrogen. Gray stick indicates a double bond. PUBLISHERS RESOURCE GROUP

HOW IT IS MADE

Nitrogen dioxide occurs naturally when nitric oxide (NO) is oxidized in the atmosphere. Nitric oxide forms naturally whenever sufficient energy is available to make possible the reaction between nitrogen and oxygen, the two primary components of the atmosphere. The formula for this reaction is



Nitric oxide, in turn, reacts readily with oxygen to form nitrogen dioxide. This reaction can be shown as



The amount of nitrogen dioxide produced naturally is so small that the gas's distinctive brown color is unnoticeable.

Such is not the case, however, in situations where nitric oxide and nitrogen dioxide are produced by human activities. The burning of coal or oil in plants that generate electricity and the combustion of gasoline in automobiles and trucks produce significant quantities of nitric oxide, which rapidly oxidizes to form nitrogen dioxide in the atmosphere. In such cases, sufficient nitrogen dioxide may be present to produce the yellowish hazy condition associated with smog.

Commercially, nitrogen dioxide is made by one of two processes, the decomposition of nitric acid (HNO_3) or the

Interesting Facts

- Scientists often use the formula NO_x to refer to the oxides of nitrogen that occur in polluted air. The two most important of those oxides are nitric oxide (NO) and nitrogen dioxide.

oxidation of ammonia (NH₃) gas. The gas can also be produced on a smaller scale or in the laboratory by a number of methods that involve the decomposition of the nitrate ion (NO₃⁻). For example, heating lead(II) nitrate [Pb(NO₃)₂] results in the formation of lead(II) oxide (PbO), nitrogen dioxide, and oxygen.

COMMON USES AND POTENTIAL HAZARDS

The most important uses of nitrogen dioxide are in the manufacture of nitric and sulfuric acids, two of the most widely used inorganic acids. The compound is also used widely as an oxidizing agent and nitrating agent. An oxidizing agent is a substance that makes oxygen available to other compounds in order to bring about a particular reaction. A nitrating agent is one that provides the nitro group (-NO₂) in producing a new compound. As an example, adding nitrate groups to the organic compound known as toluene (C₆H₅CH₃) converts it into trinitrotoluene (TNT), a powerful explosive. Nitrogen dioxide is also used in rockets, where it supplies the oxygen needed to burn the rocket fuel; as a catalyst in a number of industrial operations; in the bleaching of flour; and as a polymerization inhibitor in the manufacture of some plastics. A polymerization inhibitor is a substance that stops the formation of a polymer at some desired point in its production or use.

Nitrogen dioxide poses both safety and health hazards. As a strong oxidizing agent, it reacts readily with combustible materials, such as paper, cloth, and other organic matter to produce fires or explosions. It is also a toxic material, producing some biological effects at relatively low concentrations in the air. These effects include irritation of the eyes, nose, and throat; coughing; congestion; chest pain; and breathing

Words to Know

DIMER A molecule that consists of two identical molecules combined with each other.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

POLYMERIZATION INHIBITOR A substance that stops the formation of a polymer at some desired point in its production or use.

difficulties. The gas is sometimes referred to as an insidious agent because its effects may go unnoticed for several hours or days, during which time more serious damage may have occurred. This damage may include pulmonary edema, a condition in which the lungs begin to fill with fluid; cyanosis, a condition in which the lips and mucous membranes turn blue because of lack of oxygen; and a variety of heart problems. Long-term exposure to nitrogen dioxide may result in chronic health problems, such as hemorrhaging (blood loss), lung damage, emphysema, chronic bronchitis, and eventually death.

FOR FURTHER INFORMATION

"Cheminfo: Chemical Profiles Created by CCOHS."

<http://www.intox.org/databank/documents/chemical/nitrodix/cie748.htm> (accessed on October 20, 2005).

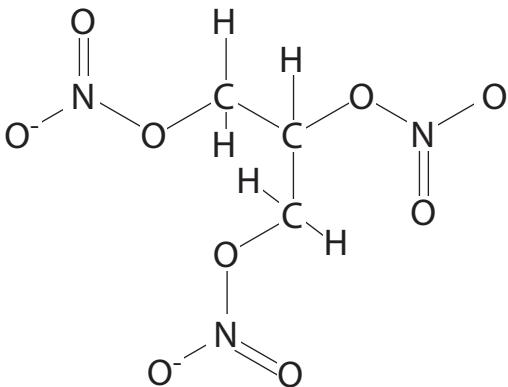
Holgate, S. T., et al. *Air Pollution and Health*. New York: Academic Press, 1999.

"NIOSH Pocket Guide to Chemical Hazards: Nitrogen Dioxide." Occupational Health & Safety Administration.
<http://www.cdc.gov/niosh/npg/npgdo454.html> (accessed on October 20, 2005).

"Nitrogen Oxides." International Programme on Chemical Safety.
<http://www.inchem.org/documents/ehc/ehc/ehc188.htm#SubSectionNumber:2.1.1> (accessed on October 20, 2005).

Patnaik, Pradyot. *Handbook of Inorganic Chemicals*. New York: McGraw Hill, 2003, 648 651.

See Also Nitric Acid



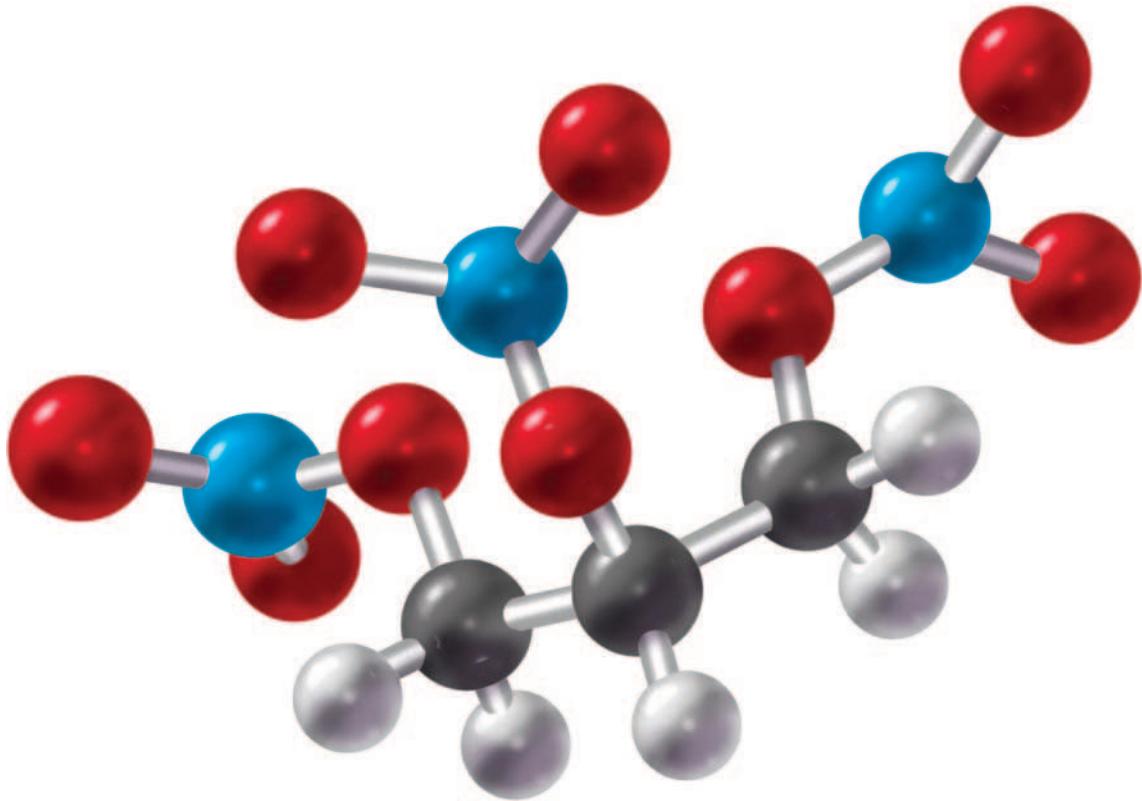
OTHER NAMES:
Trinitroglycerol; trinitroglycerin; glyceryl trinitrate
FORMULA:
$\text{CH}_2\text{NO}_2\text{CHNOCH}_2\text{NO}_2$
ELEMENTS:
Carbon, hydrogen, nitrogen, oxygen
COMPOUND TYPE:
Ester (organic)
STATE:
Liquid
MOLECULAR WEIGHT:
227.09 g/mol
MELTING POINT:
13.5°C (56.3°F)
BOILING POINT:
Explodes at 218°C 424°F
SOLUBILITY:
Slightly soluble in water; soluble in ethyl alcohol, acetone, and benzene

KEY FACTS

OVERVIEW

Nitroglycerin (nye-tro-GLIH-er-in) is a pale yellow oily flammable liquid that is highly explosive. It is used primarily as an explosive by itself and as an ingredient in dynamite. Nitroglycerin also finds application in medicine as a vasodilator, a substance that causes blood vessels to relax and open up, allowing blood to flow more freely through them.

Nitroglycerin was first developed in 1847 by the Italian chemist Ascanio Sobrero (1812–1888). Sobrero used a method of synthesis that is still the primary means of producing nitroglycerin today. He added nitric acid to glycerol, then and now a popular skin lotion, with a small amount of concentrated sulfuric acid as catalyst. He initially called his discovery *pyroglycerin*. When he placed a trace amount of nitroglycerin on his tongue, Sobrero discovered that it “gives rise to a most pulsating, violent headache, accompanied by a great weakness of the limbs.” He also discovered that the chemical was highly combustible and explosive. He considered the substance so dangerous that he warned against its



Nitroglycerin. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atoms are nitrogen. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

use and did not disclose his discovery to the world for more than a year.

In the 1860s, the Swedish chemist Alfred Nobel (1833–1896) developed a process for manufacturing nitroglycerin on a large scale. His company sold a combination of nitroglycerin and gunpowder called Swedish blasting oil, but the product proved far too dangerous to use. Several accidents involving the substance occurred. One explosion in Nobel's factory killed several people, including his brother Emil. To make a safer explosive, Nobel found a way to combine nitroglycerin with clay, a chemically inactive material. He called the combination dynamite. Dynamite soon became the explosive of choice in construction, demolition, and mining projects around the world.

Nitroglycerin's medical uses were first explored in detail by Sir Thomas Lauder Brunton (1844–1916) of the Royal Infirmary in Edinburgh, Scotland. In 1938, the U.S. Food

Interesting Facts

- Before Nobel began studying the explosive properties of nitroglycerin, people reportedly used the compound for everyday chores, such as polishing boots and greasing wagon wheels, and even as lamp oil. The consequences were often fatal.
- Although he spent most of his life developing and manufacturing explosives, Nobel was a humanitarian who wanted to see technology used for the benefit of society and the advancement of world

peace. His handwritten will, although fiercely contested, provided for the creation of the Nobel Foundation. The foundation's primary responsibility is to award prizes in chemistry, physics, physiology or medicine, and peace every year. The first Nobel prizes were awarded in 1901. Today, a Nobel prize is regarded as the highest honor given in science. Each prize is worth over a million dollars in cash.

and Drug Administration first approved the use of nitroglycerin as a vasodilator for the treatment of heart problems.

HOW IT IS MADE

The preparation of nitroglycerin is a straightforward, but very delicate, chemical procedure. When nitric acid is added to glycerol, three nitrate groups (-NO₂) replace each of the three hydroxyl (-OH) groups on glycerol. This reaction occurs only if the acid and glycerol are moderately warm, at least at room temperature. But combining nitric acid, glycerol, and sulfuric acid (the catalyst) for the reaction generates heat. Too much heat causes the nitroglycerin being produced to explode. The resolution for this dilemma is to begin the reaction at room temperature, but then encase the reaction vessel in ice as soon as the reaction begins. The ice absorbs heat generated during the reaction, allowing the formation of nitroglycerin without it becoming too warm.

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

COMMON USES AND POTENTIAL HAZARDS

Nitroglycerin's primary use is as an explosive, either by itself or as a component of dynamite. Today, it is marketed under more than 60 trade names, including Coro-Nitro®, Deponit®, GTN®, Nitroglin®, Nitrong®, Perglottal®, Reminitrol®, Sustac®, Tridil®, and Vasoglyn®. Nitroglycerin is classified as a high explosive, which means that it explodes very rapidly with a very large force. It is detonated ("set off") either by heat or by shock. Because of its very unstable character, it is usually transported at low temperatures (5°C to 10°C ; 40°F to 50°F), at which it is more stable.

When used as a heart medication, nitroglycerin is administered in the form of a pill, patch, or intravenous solution. Nitroglycerin works in the body by releasing nitric oxide, a vasodilator. Vasodilators cause the smooth muscle surrounding blood vessels to relax, allowing the vessels to expand and improve the flow of blood. The drug is often taken when the pain of angina or a heart attack is first noticed.

People who come into contact with nitroglycerin in the workplace are at risk for a number of health problems. The compound is a skin, eye, and respiratory system irritant. It may cause nausea, vomiting, abdominal cramps, headache, mental confusion, delirium, sweating, a burning sensation on the tongue, paralysis, convulsions, and death.

FOR FURTHER INFORMATION

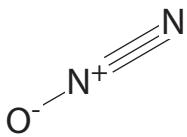
"Nitroglycerin." Imperial College of London, Department of Chemistry.
http://www.ch.ic.ac.uk/rzepa/mim/environmental/html/nitroglyc_text.htm (accessed on October 20, 2005).

"Nitroglycerin: Dynamite for the Heart." *Chemistry Review* (November 1999): 28.

Rawls, Rebecca. "Nitroglycerin Explained." *Chemical & Engineering News* (June 10, 2002): 12.

"Why Is Nitroglycerin Explosive?" General Chemistry Online
<http://antoine.frostburg.edu/chem/senese/101/consumer/faq/nitroglycerin.shtml> (accessed on October 20, 2005).

See Also Nitric Acid; Sulfuric Acid

**OTHER NAMES:**

See Overview.

FORMULA:

N₂O

ELEMENTS:

Nitrogen; oxygen

COMPOUND TYPE:

Nonmetallic oxide
(inorganic)

STATE:

Gas

MOLECULAR WEIGHT:

44.01 g/mol

MELTING POINT:

-90.8°C (-131°F)

BOILING POINT:

-88.48°C (-121.3°F)

SOLUBILITY:

Slightly soluble in water; soluble in ethyl alcohol and ether

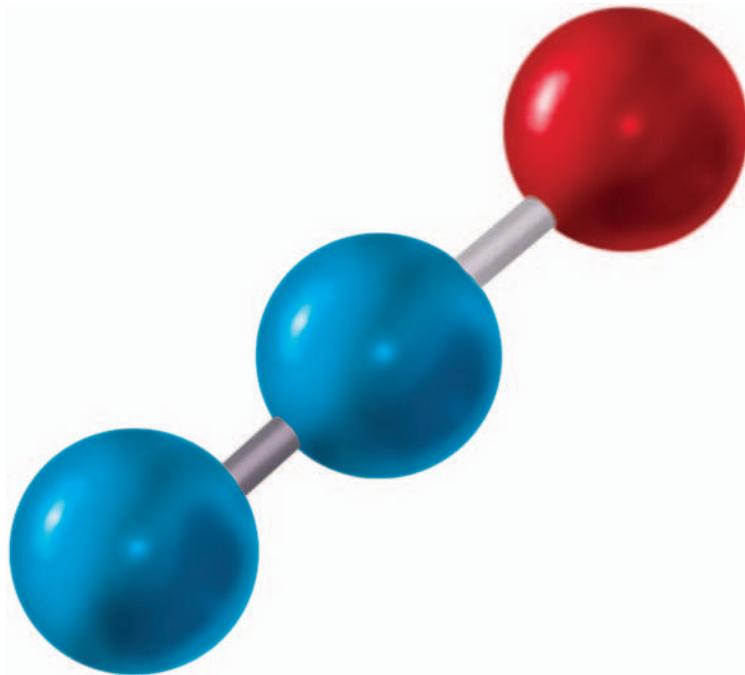
K E Y F A C T S**OVERVIEW**

Nitrous oxide (NYE-truss OX-side) is also known as dinitrogen oxide, dinitrogen monoxide, nitrogen monoxide, and laughing gas. It is a colorless, nonflammable gas with a sweet odor. Its common name of laughing gas is derived from the fact that it produces a sense of light-headedness when inhaled. The gas is widely used as anesthetic, a substance that reduces sensitivity to pain and discomfort.

Nitrous oxide was probably first produced by the English chemist and physicist Robert Boyle (1627-1691), although he did not recognize the new compound he had found. Credit for the discovery of nitrous oxide is, therefore, usually given to the English chemist Joseph Priestley (1733-1804), who produced the gas in 1772 and named it "nitrous air." Other early names used for the gas include "gaseous of azote" (nitrogen) and "oxide of speton." The most complete experiments on the gas were conducted by the English chemist and physicist Sir Humphry Davy (1778-1829), who tested nitrous oxide on himself and his friends. He found that the gas could lessen

Nitrous oxide. Red atom is oxygen and blue atoms are nitrogen. The nitrogen atoms share a triple bond.

PUBLISHERS RESOURCE GROUP



pain and discomfort and provided a sense of relaxation and well-being. Before long, doctors were making use of Davy's discovery by using nitrous oxide as an anesthetic.

The public found other uses for the gas as well. During the Victorian period in England, members of the upper class often held laughing gas parties at which people gathered to inhale nitrous oxide as a recreational drug, rather than for any therapeutic purpose. In the United States, the showman P. T. Barnum (1810–1891) created a sideshow exhibit in which people were invited to test the effects of inhaling nitrous oxide. After seeing a demonstration of this kind, the American dentist Horace Wells (1815–1848) first used nitrous oxide as an anesthetic on his patients.

In 1868, the American surgeon Edmund Andrews (1824–1904) extended the use of nitrous oxide as an anesthetic for his surgical patients. He mixed the gas with oxygen to ensure that patients received enough oxygen while receiving the anesthetic. The gas is still widely used by dentists as a safe and relatively pleasant way of helping patients endure the discomfort of drilling and other dental procedures.

Interesting Facts

- Humphry Davy proposed the name laughing gas for nitrous oxide.
- In the United Kingdom, nitrous oxide is often used as an anesthetic for women about to give birth.
- In the 1830s, Samuel Colt (1814–1862), inventor of the Colt 45 revolver, toured North America, giving laughing gas demonstrations.

HOW IT IS MADE

The most common commercial method of producing nitrous oxide involves the controlled heating of ammonium nitrate (NH_4NO_3). The compound decomposes to form nitrous oxide and water. The reaction is essentially the same one originally used by Priestley in 1772. Although an efficient means of producing the gas, the reaction must be carried out with extreme care as ammonium nitrate has a tendency to decompose explosively when heated. Nitrous oxide can also be produced by the decomposition of nitrates (compounds containing the NO_3^- radical), nitrites (compounds containing the NO_2^- radical), or nitriles (compounds containing the CH^- radical).

COMMON USES AND POTENTIAL HAZARDS

Nitrous oxide is best known and most widely used as an anesthetic. Its use is limited primarily to dental procedures and minor surgeries. Dentists favor nitrous oxide as an anesthetic because the gas does not make patients completely unconscious and does not require an anesthesiologist to administer it. Nitrous oxide works as an anesthetic by blocking neurotransmitter receptors in the brain, preventing pain messages from being transmitted.

Nitrous oxide is also used as a fuel additive in racing cars, in which case it is often referred to as nitro. The gas is injected into the intake manifold where it mixes with air and fuel vapors. Since it breaks down at the high temperatures in the car's engine, it provides additional oxygen to increase the

Words to Know

ANESTHETIC A substance that reduces sensitivity to pain and discomfort.

NITRATE A compound that includes the radical consisting of one nitrogen atom and three oxygen atoms (NO_3).

NITRITE A compound that includes the radical consisting of one nitrogen atom and two oxygen atoms (NO_2).

RADICAL A group of atoms bonded together that act like a single entity in chemical reactions.

efficiency with which the fuel burns. During World War II, pilots used nitrous oxide for a similar purpose in their airplanes.

Some additional uses of nitrous oxide include:

- As a propellant in food aerosols;
- For the detection of leaks;
- As a packaging gas for potato chips and other snack foods, preventing moisture from making the product become stale;
- In the preparation of other nitrogen compounds; and
- As an oxidizing agent for various industrial processes.

Nitrous oxide is safe to use in moderate amounts under controlled conditions. Some people use the compound as a recreational drug, however, hoping to get a “high” from inhaling it. One risk of this practice is that the inhalation of nitrous oxide may reduce the amount of oxygen a person receives. Also, some long-term health effects, such as anemia (low red blood cell count) and neuropathy (damage to the nerves), have been associated with excessive use of the compound. The use of nitrous oxide for recreational purposes is a crime in some states.

FOR FURTHER INFORMATION

“Gas Data: Nitrous Oxide.” Air Liquide.

<http://www.airliquide.com/en/business/products/gases/gasdata/index.asp?GasID=55> (accessed on October 20, 2005).

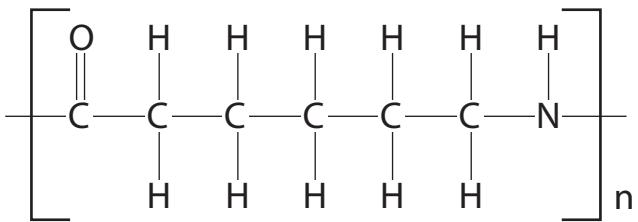
Neff, Natalie. ‘No Laughing Matter.’ *Auto Week* (May 19, 2003): 30.

"Nitrogen Oxide." Center for Advanced Microstructures and Devices, Louisiana State University.
http://www.camd.lsu.edu/msds/n/nitrous_oxide.htm#Synonyms
(accessed on October 20, 2005).

"Occupational Safety and Health Guideline for Nitrous Oxide." Occupational Safety and Health Administration.
<http://www.osha.gov/SLTC/healthguidelines/nitrousoxide/recognition.html> (accessed on October 20, 2005).

Pae, Peter. "Sobering Side of Laughing Gas." *Washington Post* (September 16, 1994): B1.

See Also Ammonium Nitrate



FORMULA:

Nylon 6:
 $-\text{[CO(CH}_2\text{)}_5\text{NH-]}_n$;
 Nylon 66:
 $-\text{[CO(CH}_2\text{)}_4\text{CO-NH(CH}_2\text{)}_6\text{NH-]}_n$

ELEMENTS:

Carbon, hydrogen,
oxygen, nitrogen

COMPOUND TYPE:

Polymer (organic)

STATE:

Solid

MOLECULAR WEIGHT:

Very large

MELTING POINT:

Nylon 6: 223°C
 (433°F); Nylon 66:
 265°C (509°C)

BOILING POINT:

Decomposes above
 melting point

SOLUBILITY:

Insoluble in water and
 most organic
 solvents; soluble in
 strong acids

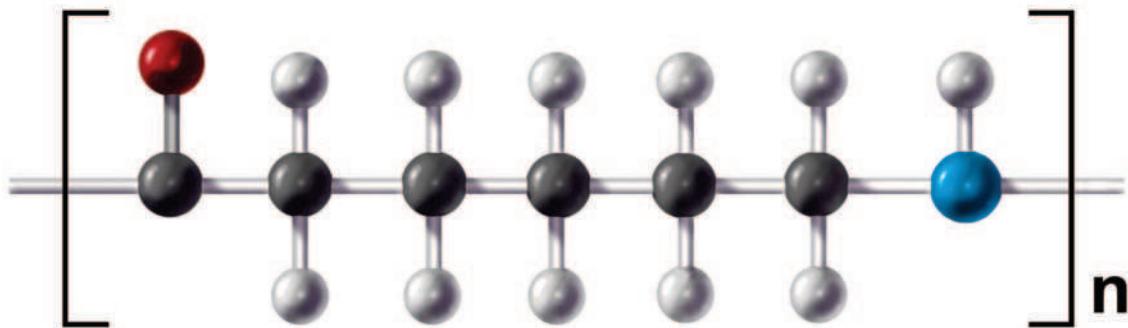
K E Y F A C T S

OVERVIEW

The term nylon is used to describe a family of organic polymers called the polyamides that contain the amide (-CONH) group. The members of the family are distinguished from each other by a numbering system indicating the chemical composition of the polymer molecule. The two most important nylons are nylon 6 and nylon 66 which, between them, account for nearly all of the nylon produced in the United States. Other nylons that are produced in much smaller amounts include nylon 11, nylon 12, nylon 46, and nylon 612.

The polyamides are thermoplastic polymers. The term "thermoplastic" means that the polymer can be repeatedly melted and hardened by alternate heating and cooling. By contrast, certain other types of polymers, known as thermo-setting plastics, can not be re-melted once they have hardened.

Nylon was invented in 1935 by Wallace Carothers (1896–1937), an employee of the DuPont Chemical Corporation at the time. Carothers was searching for a synthetic substitute



Nylon 6. Red atom is oxygen;
white atoms are hydrogen;
black atoms are carbon; blue
atoms are nitrogen; gray stick
shows a single bond.

PUBLISHERS RESOURCE GROUP

of silk because natural supplies were insufficient to meet growing demands for that fiber. Using coal, water, and air as raw materials, Carothers developed a synthetic product that could be stretched into a fiber and that got stronger and silkier as it was stretched. Nylon first appeared in commercial products in 1938, first in toothbrushes with nylon bristles, and later in women's stockings. During World War II, the U.S. military found a number of uses for the compound, especially the manufacture of parachutes. Today, nylon is used in a very wide array of products.

HOW IT IS MADE

Nylon 66 is still made commercially by the procedure originally developed by Carothers in 1935. The process is initiated with the reaction between adipic acid ($\text{HOOC}(\text{CH}_2)_4\text{COOH}$) and hexamethylenediamine ($\text{H}_2\text{N}(\text{CH}_2)_6\text{NH}_2$), which results in the formation of the monomer ($\text{-[CO}(\text{CH}_2)_4\text{CONHCH}_2\text{]}_6\text{-[COCH}_2\text{]}_4\text{CONH-}$) from which the polymer develops. That monomer has active groups at both ends of the molecule, allowing it to react with others of its kind to form long chains that make up the polymer. Methods for preparing the other polyamides can differ quite significantly from the Carothers procedure. For example, nylon 6 is made by the polymerization of aminocaproic acid ($\text{H}_2\text{N}(\text{CH}_2)_5\text{COOH}$). The polymer formed in this reaction also has blocks of methylene ($-\text{CH}_2-$) groups joined by amide (CONH) groups, although the specific structure is somewhat different than it is for nylon 66.

COMMON USES AND POTENTIAL HAZARDS

Production of nylon 6 and nylon 66 for 2006 in the United States was estimated to be about 1.6 million kilograms (3.5 mil-

Interesting Facts

- The term nylon has never been trademarked. The name itself has no specific meaning. The -on suffix was chosen, however, to suggest a comparison with other fibers, such as cotton and rayon.
- DuPont originally thought of calling nylon No-Run. They abandoned that plan, however, when they realized that nylon stockings do, in fact, run.

lion pounds). About three-quarters of that production is used for the manufacture of fibers used in industrial operations and for textiles. Fabrics made from nylon are strong, resistant to abrasion, resistant to alkaline materials, and capable of taking most dyes. The clothing industry uses nylon fibers to make a variety of fabric types, including fleece, velvet, satin, taffeta, lace, and seersucker. Nylon clothing is lightweight, silky, attractive and smooth. Since it does not absorb moisture or odors readily, it is in demand for the production of athletic clothing. Nylon is also used in the manufacture of home textiles, such as furniture covering. Nylon rope is twice as strong as rope made from natural fibers such as hemp and weighs less.

About 15 to 20 percent of all nylon produced in the United States goes to the manufacture of molded plastics, such as those used in automobile and truck parts, housing for electrical devices, and consumer articles. The next largest application for nylon is in tubing, pipes, films, and coatings. Some specific articles made from nylon 6 and nylon 66 include the following:

- Reinforcing cord for tires;
- Fuel tanks for automobiles;
- Tennis rackets;
- Fishing nets and lines;
- Towlines for gliders;
- Women's stockings;
- Sutures for surgical procedures;

Words to Know

POLYAMIDE An organic polymer that contains the amide (-CONH) group.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

THERMOPLASTIC Capable of being heated so that it can be reshaped, then cooled so that it hardens.

THERMOSETTING Capable of being heated so that it hardens.

- Bristles for toothbrushes, hairbrushes, and paint brushes;
- Parachutes;
- Pen tips; and
- Artificial turf for athletic fields.

One of the most important uses for nylon historically was the manufacture of women's stockings. Silk stockings were an essential item of a well-dressed woman's ensemble in the 1920s and 1930s. They made a woman's legs look smooth and sleek, and they were the ideal accompaniment to short skirts and high heels. In the 1920s, women spent millions of dollars on silk stockings. The DuPont company began the search for a silk substitute because it hoped to develop a product that could be used to make stockings that looked and felt as if they were made of silk, but that cost less. Nylon stockings became commercially available in 1939, and by 1941, some sixty million pairs had been sold. Both nylon and silk stockings were difficult to get during World War II because the fibers were being used for military applications. As soon as the war was over, however, women began buying nylon stockings in large quantities once again.

FOR FURTHER INFORMATION

Hermes, Matthew E. *Enough for One Lifetime: Wallace Carothers, Inventor of Nylon*. Philadelphia: Chemical Heritage Foundation, 1996.

"Nylons (Polyamides)." MatWeb.

<http://www.matweb.com/reference/nylon.asp> (accessed on October 24, 2005).

"Nylon Stockings." The Great Idea Finder.

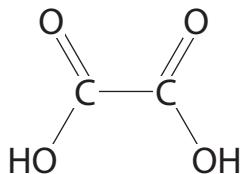
<http://www.ideafinder.com/history/inventions/story062.htm>
(accessed on October 24, 2005).

"Polyamide 6 Nylon 6 PA 6." Azom.com.

<http://www.azom.com/details.asp?ArticleID=442> (accessed
on October 24, 2005).

"Polyamides." Chemguide.

<http://www.chemguide.co.uk/organicprops/amides/polyamides.html> (accessed on October 24, 2005).



OTHER NAMES:	Ethanedioic acid
FORMULA:	<chem>HOOCOOOH</chem>
ELEMENTS:	Carbon, hydrogen, oxygen
COMPOUND TYPE:	Dicarboxylic acid; organic acid (organic)
STATE:	Solid
MOLECULAR WEIGHT:	90.04 g/mol
MELTING POINT:	189.5°C (373.1°F); begins to sublime at 157°C (315°F)
BOILING POINT:	Not applicable
SOLUBILITY:	Soluble in water and ethyl alcohol; slightly soluble in ether; insoluble in most other organic solvents

KEY FACTS

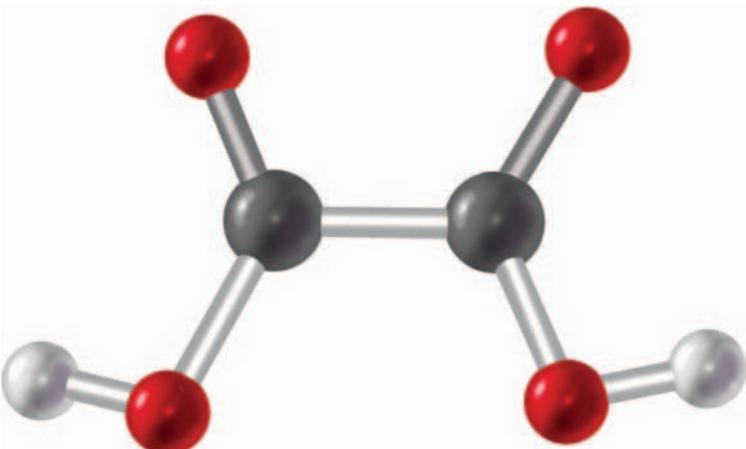
OVERVIEW

Oxalic acid (ok-SAL-ik AS-id) is a transparent, colorless, crystalline solid that often occurs as the dihydrate (HOOC-COOH·2H2O). The dihydrate melts and begins to decompose at 101.5°C (214.7°F), forming the anhydrous acid. The compound is odorless, but has a characteristic tart, acidic taste. The acid should never be tasted, however, as it is very toxic.

Oxalic acid is one of the first organic acids to have been discovered and studied. It was first isolated by the German chemist Johann Christian Wiegbleb (1732-1800) in 1769 and first synthesized by the Swedish chemist Karl Wilhelm Scheele (1742-1786) in 1776. Friedrich Wöhler's (1800-1882) synthesis of oxalic acid entirely from inorganic materials was a critical step in disproving the Vitalistic Theory of chemistry. The theory claimed that compounds found in living organisms could be produced only by the act of some supernatural being, and not by human actions.

Oxalic acid occurs naturally in a number of vegetable products, including spinach, rhubarb, tea, chocolate, oats,

Oxalic acid. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP



pumpkin, lentils, beets, parsnips, and many kinds of nuts and berries. The amount present in foods is generally so low that it does not present risk to people who eat such products. Oxalic acid also occurs as a product of carbohydrate metabolism in animals.

HOW IT IS MADE

Traditionally, oxalic acid has been extracted from natural products by treating them with an alkaline solution, followed by crystallization of the acid. Sodium hydroxide is the alkaline material most commonly used for this procedure. Today, a number of methods are available for the commercial preparation of oxalic acid. In one procedure, carbon monoxide gas is bubbled through a concentrated solution of sodium hydroxide to produce oxalic acid. Alternatively, sodium formate (COONa) is heated in the presence of sodium hydroxide or sodium carbonate to obtain the acid. Another popular method of preparing oxalic acid involves the oxidation of sucrose (common table sugar) or more complex carbohydrates using nitric acid as a catalyst. The reaction results in the formation of oxalic acid and water as the primary products.

COMMON USES AND POTENTIAL HAZARDS

Over a half million kilograms (about 1 million pounds) of oxalic acid are produced in the United States each year.

Interesting Facts

- One family of plants, the Oxalidaceae, get their name from the high concentrations of oxalic acid they contain. Perhaps the best known member of the family is the wood sorrel.
- A number of molds produce oxalic acid as a major metabolic product. Some species of *Penicillium* and *Aspergillus*, for example, convert glucose into oxalic acid.

The greatest portion of the product is used in a variety of cleaning products, including substances for the bleaching and cleaning of wood, cork, cane, feathers, and natural and synthetic fibers. Many metal polishes, auto radiator cleaners, and laundry rinses also contain oxalic acid. Some rust-proofing materials also contain the compound.

The list of other household and industrial applications of oxalic acid is extensive and includes:

- In the printing and dyeing of fabrics, especially calico;
- For the treatment of leather products, especially the leather used to make book covers;
- As a paint and varnish remover as well as a stain remover for ink and rust marks;
- In the manufacture of blue ink, celluloid, rubber, and other synthetic products;
- For the decolorization of glycerol and the stabilization of hydrocyanic acid (HCN);
- As a purification agent for various chemicals, especially methanol (methyl alcohol); and
- In the preparation of certain medicines and pharmaceuticals.

Oxalic acid is a strong skin, eye, and respiratory irritant in pure form. It is also toxic by ingestion, causing nausea, vomiting, diarrhea, kidney damage, convulsions, coma, and

Words to Know

ANHYDROUS COMPOUND A compound that lacks any water of hydration.

HYDRATE A chemical compound formed when one or more molecules of water is added physically to the molecule of some other substance.

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

METABOLISM Process that includes all of the chemical reactions that occur in cells

by which fats, carbohydrates, and other compounds are broken down to produce energy and the compounds needed to build new cells and tissues.

SUBLIME To go from solid to gaseous form without passing through a liquid phase.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

death. Since people are exposed to pure oxalic acid only in the workplace, these hazards are usually not of concern to most individuals. Of more concern to the general public is the possibility of ingesting unusually large amounts of foods containing oxalic acid. In the body, the acid tends to react with calcium ions (Ca^{2+}) forming calcium oxalate, which is insoluble. This reaction has two harmful effects. First, it removes calcium needed by the body for other biological functions, resulting in a calcium deficiency problem. Second, it may result in the precipitation of calcium oxalate crystals on the inner lining of blood vessels and the small tubes in the kidneys, reducing the flow of blood and resulting in the development of kidney disorders.

FOR FURTHER INFORMATION

“Material Safety Data Sheet.” Hill Brothers Chemical Company.
http://hillbrothers.com/msds/pdf/oxalic_acid.pdf (accessed on October 23, 2005).

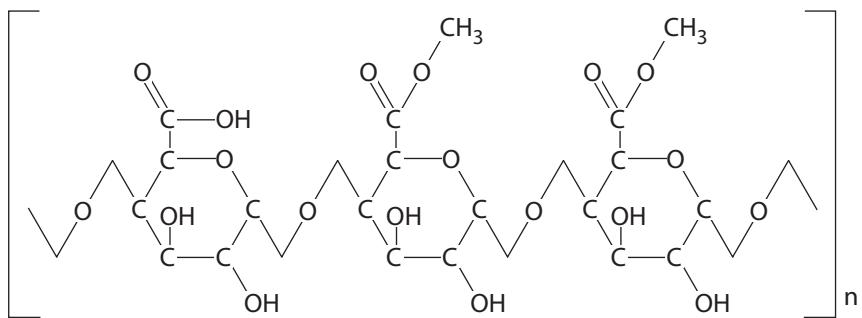
“Oxalic Acid.” Al’s Home Improvement Center.
http://alsnetbiz.com/homeimprovement/oxalic_acid.html (accessed on October 23, 2005).

"Oxalic Acid Dihydrate." Chemical Land 21.

<http://www.chemicalland21.com/arokorhi/industrialchem/organic/OXALIC%20ACID.htm> (accessed on October 23, 2005).

"The Rhubarb Compendium." Rhubarb Poison Information Center.

<http://www.rhubarbinfo.com/rhubarb poison.html> (accessed on October 23, 2005).



KEY FACTS	
FORMULA:	Not applicable
ELEMENTS:	Carbon, hydrogen, oxygen, and other elements
COMPOUND TYPE:	Not applicable
STATE:	Solid
MOLECULAR WEIGHT:	Varies widely: 20,000 to 400,000 g/mol
MELTING POINT:	Not applicable
BOILING POINT:	Not applicable
SOLUBILITY:	Soluble in water; insoluble in organic solvents

Pectin

OVERVIEW

Pectin (PEK-tin) is a mixture, not a compound. Mixtures differ from compounds in a number of important ways. The parts making up a mixture are not chemically combined with each other, as they are in a compound. Also, mixtures have no definite composition, but consist of varying amounts of the substances from which they are formed.

Chemically, pectin is a polysaccharide, a very large molecule made of many thousands of monosaccharide units joined to each other in long, complex chains. Monosaccharides are simple sugars. The most familiar monosaccharide is probably glucose, the sugar from which the human body obtains the energy it needs to grow and stay healthy. The monosaccharides in pectin are different from and more complex than glucose.

Pectin occurs naturally in many fruits and vegetables. It is most abundant in citrus fruits such as lemons, oranges, and grapefruits, which may consist of up to 30 percent pectin. In pure form it is a yellowish-white powder with virtually no

odor and a slightly gummy taste. When dissolved in water, it forms a thick, jelly-like mass. This property explains one of its primary purposes: the jelling of fruits when they are made into jams and jellies.

HOW IT IS MADE

Pectin is made naturally in ripening fruit. It is obtained commercially by treating the raw material (citrus peel or apple pomace) with hot, acidified water. (Apple pomace is the residue remaining after pressing of apples.) The pectin in the peel or apple pomace dissolves in the hot water and is then purified by repeated filtrations. It is extracted from the water solution by adding alcohol or an aluminum salt to the solution, causing the pectin to precipitate out of solution. The precipitate is then dried and ground into a powder.

Additional steps are sometimes carried out to convert the pectin produced by this method, called high ester pectin, to a form that is more soluble: low ester pectin. To achieve this change, high ester pectin is treated with either acids or alkalis, washed, and purified.

COMMON USES AND POTENTIAL HAZARDS

Pectin is used primarily as a jelling agent in the manufacture of jams and jellies. It also has a number of other applications as a food additive. For example, it is added to some yogurts to provide the consistency that allows the yogurt to hold its shape and still be capable of being stirred. It is added to concentrated fruit drinks to keep the solid and liquid components of the drink in suspension with each other. It is also an ingredient in fruit and milk desserts, added to ensure that the final product has the proper consistency and stability.

Pectin is also used as an additive in pharmaceutical and cosmetic preparations. It acts as an emulsifying agent, to stabilize the product, or to give it the proper consistency. In combination with an antibiotic, pectin has also been used as an anti-diarrheal agent. Some studies have shown that daily doses of pectin may have a small but significant lowering effect on cholesterol levels.

The U.S. Food and Drug Administration has classified pectin as an approved food additive. It is considered safe

Interesting Facts

- The role of pectin in the formation of jams and jellies from fruits was first recognized in the 1820s.
- The first commercial manufacture of pectin took place in Germany in 1908. Producers of apple juice found that they could use apple pomace, previously a waste product, to make a useful product that could be sold, pectin.
- The first recipes for the use of pectin to make jams and jellies date to the first century when the Roman writer Marcus Gavius Apicius wrote a recipe book "Of Culinary Methods.";
- American inventor Paul Welch was granted a patent in 1917 for the production of grape jam, using pectin. Welch called his product Grapelade and sold his entire production to the U.S. Army. The Army sent the Grapelade to troops serving in Europe in World War I (1914–1918). After the war, returning troops demanded more Grapelade, making it a popular consumer item in the United States.

for human consumption when used in normal amounts as a food additive. It may cause some digestive problems for people with allergies to citrus fruits. Some studies suggest that pectin may also inhibit the absorption of minerals such as zinc, copper, iron, and calcium, although this effect is not serious enough to prevent its use as a food additive.

Words to Know

EMULSION A temporary mixture of two liquids that normally do not dissolve in each other.

POLYSACCHARIDE A very large molecule made of many thousands of simple sugar molecules joined to each other in long, complex chains.

FOR FURTHER INFORMATION

“Genu® Pectin.” CP Kelco.

<http://www.cpkelco.com/food/pectin.html> (accessed on December 22, 2005).

Knox, J. Paul, and Graham B. Seymour, eds. *Pectins and Their Manipulation*. Boca Raton, FL: CFC Press, June 2002.

“Pectin.”

http://www.cfs.purdue.edu/class/f&n63o/Virt_Class_2/pectin.htm (accessed on December 22, 2005).

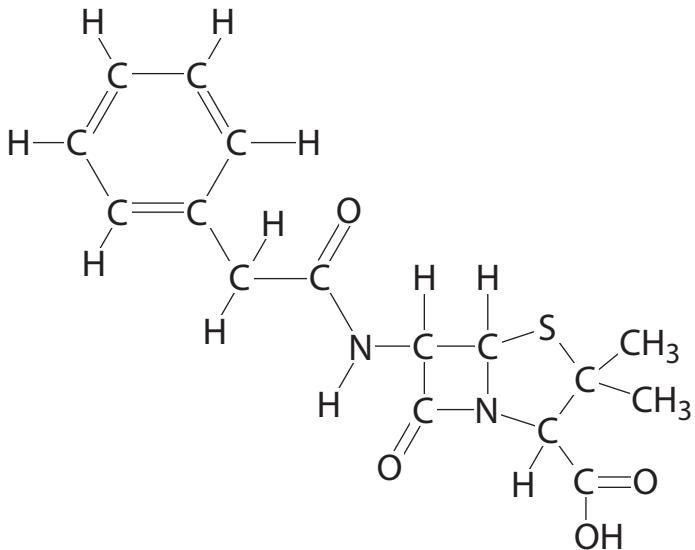
“Pectin.” PDRHealth.

http://www.pdrhealth.com/drug_info/nmdrugprofiles/nutsupdrugs/pec_0198.shtml (accessed on December 22, 2005).

See Also Gelatin

OTHER NAMES:	Not applicable; see Overview
FORMULA:	$(CH_3)_2C_5H_3NSO$ $(COOH)NHCOR$, where R represents any one of a number of substituted groups; see Overview
ELEMENTS:	Carbon, hydrogen, oxygen, nitrogen, sulfur
COMPOUND TYPE:	Bicyclic acid (organic)
STATE:	Solid
MOLECULAR WEIGHT:	Varies; see Overview g/mol
MELTING POINT:	Varies; see Overview
BOILING POINT:	Not applicable; all forms decompose when heated above their melting points
SOLUBILITY:	Slightly soluble in water; soluble in ethyl alcohol, ether, chloroform and most organic solvents

KEY FACTS

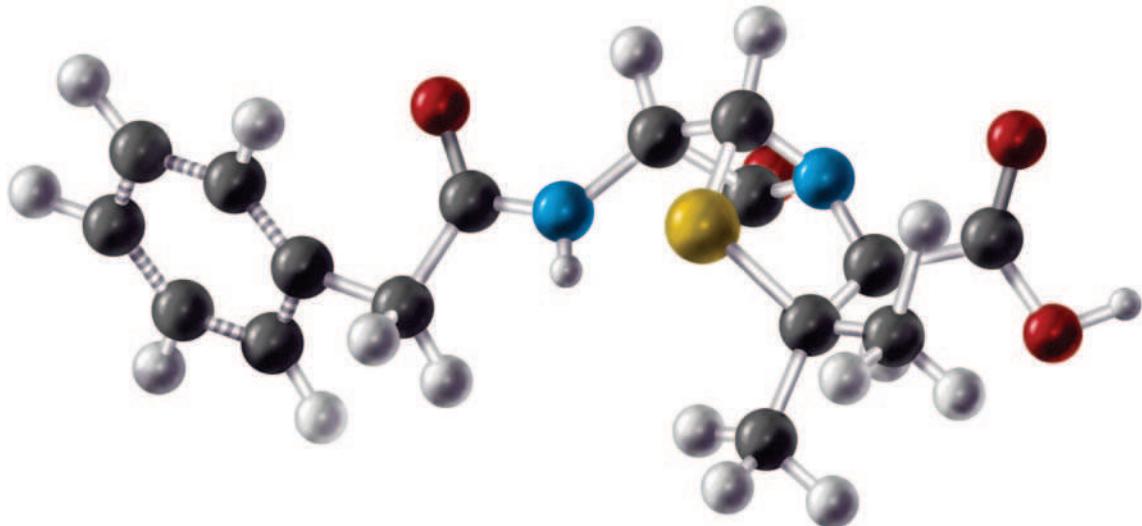


Penicillin

OVERVIEW

The penicillins (pen-uh-SILL-ins) are a class of antibiotic compounds derived from the molds *Penicillium notatum* and *Penicillium chrysogenum*. The class contains a number of compounds with the same basic bicyclic structure to which are attached different side chains. That basic structure consists of two amino acids, cysteine and valine, joined to each other to make a bicyclic (“two-ring”) compound. The different forms of penicillin are distinguished from each other by adding a single capital letter to their names. Thus: penicillin F, penicillin G, penicillin K, penicillin N, penicillin O, penicillin S, penicillin V, and penicillin X. A number of other antibiotics, including ampicillin, amoxicillin, and methicillin, have similar chemical structures.

Penicillin was discovered accidentally in 1928 by the Scottish bacteriologist Alexander Fleming (1881-1955). Fleming noticed that a green mold, which he later identified as *Penicillium notatum*, had started to grow on a petri dish that he had coated with bacteria. As the bacteria grew towards the



Penicillin. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; blue atoms are nitrogen; and yellow atom is sulfur. Gray sticks show double bonds; striped sticks indicate a benzene ring. PUBLISHERS RESOURCE GROUP

mold, they began to die. At first, Fleming saw some promise in this observation. Perhaps the mold could be used to kill the bacteria that cause human disease. His experiments showed, however, that the mold's potency declined after a short period of time. He was also unable to isolate the anti-bacterial chemical produced by the mold. He decided that further research on *Penicillium* was probably not worthwhile.

As a result, it was not until a decade later that *Penicillium*'s promise was realized. In 1935, English pathologist Howard Florey (1898-1968) and his biochemist colleague Ernst Chain (1906-1970) came across Fleming's description of his experiment and decided to see if they could isolate the chemical product produced by *Penicillium* with anti-bacterial action. They were eventually successful, isolating and purifying a compound with anti-bacterial action, and, in 1941, began trials with human subjects to test its safety and efficacy (ability to kill bacteria). The successful conclusion of those trials not only provided one of the great breakthroughs in the human battle against infectious diseases, but also won for Florey, Chain, and Fleming the 1945 Nobel prize for Physiology or Medicine.

HOW IT IS MADE

Penicillins are classified as biosynthetic or semi-synthetic. Biosynthetic penicillin is natural penicillin. It is

Interesting Facts

- The discovery of penicillin's antibacterial action came at just the right time: the onset of World War II. The new drug made possible the saving of untold numbers of lives. As just one example, about three-quarters of all soldiers who developed bone infections as a result of wounds suffered in

World War I died of those infections. By contrast, no more than about 5 percent of those who developed similar infections in World War II died. The availability of penicillin to treat those wounds in World War II made the difference in those survival rates.

produced by culturing molds in large vats and collecting and purifying the penicillins they produce naturally. There are six naturally occurring penicillins. The specific form of penicillin produced in a culturing vat depends on the nutrients provided to the molds. Of the six natural penicillins, only penicillin G (benzylpenicillin) is still used to any extent.

Semi-synthetic penicillins are produced by making chemical alterations in the structure of a naturally occurring penicillin. For example, penicillin V is made by replacing the $\text{-CH}_2\text{C}_6\text{H}_5$ group in natural penicillin G with a $\text{-CH}_2\text{OC}_6\text{H}_5$ group.

COMMON USES AND POTENTIAL HAZARDS

Penicillins are prescription medications used to treat a variety of bacterial infections, including meningitis, syphilis, sore throats, and ear aches. They do so by inactivating an enzyme used in the formation of bacterial cell walls. With the enzyme inactivated, bacteria can not make cell walls and die off. Penicillins do not act on viruses in the same way they do on bacteria, so they are not effective against viral diseases, such as the flu or the common cold.

A number of side effects are related to the use of penicillin. These side effects include diarrhea, upset stomach, and vaginal yeast infections. In those individuals who are allergic

Words to Know

BICYCLIC Refers to a molecule in which the atoms are arranged so as to look like two rings.

BIOSYNTHETIC Natural; made from a living organism.

SEMI-SYNTETIC Produced by making chemical alterations in the structure of a naturally occurring organism.

to penicillins, side effects are far more serious and include rash, hives, swelling of tissues, breathing problems, and anaphylactic shock, a life-threatening condition that requires immediate medical treatment.

Penicillin may alter the results of some medical tests, such as those for the presence of sugar in the urine. Penicillin can also interact with a number of other medications, including blood thinners, thyroid drugs, blood pressure drugs, birth control pills, and other antibiotics, in some cases decreasing their effectiveness.

Once promoted as wonder drugs, the use of penicillins has declined slowly because of the spread of antibiotic resistance. Antibiotic resistance occurs when new strains of bacteria evolve that are resistant to existing types of penicillin. One reason that antibiotic resistance has become a problem is the extensive and often unnecessary use of penicillins. When they are prescribed for colds and the flu, for example, they have no effect on the viruses that cause those diseases, but they encourage the growth of bacteria more able to survive against penicillins.

FOR FURTHER INFORMATION

“ β Lactam Antibiotics: Penicillins.” *The Merck Manual of Diagnosis and Therapy*. Chapter 153. Available online at <http://www.merck.com/mrkshared/mmanual/section13/chapter153/153b.jsp> (accessed on October 23, 2005).

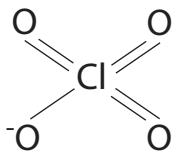
Chain, E. B. “The Chemical Structure of the Penicillins.” Nobel Lecture, March 20, 1945. Available online at <http://nobelprize.org/medicine/laureates/1945/chain lecture.pdf> (accessed on October 23, 2005).

Moore, Greogry A., and Ollie Nygren. "Penicillins." The Nordic Expert Group for Criteria Documentation of Health Risks from Chemicals.

http://ebib.arbetslivsinstitutet.se/ah/2004/ah2004_06.pdf
(accessed on October 23, 2005).

Ross Flanigan, Nancy. "Penicillins." *Gale Encyclopedia of Medicine*. Edited by Jacqueline L. Longe and Deidre S. Blanchfield. 2nd ed. Detroit, MI: Gale, 2002.

"Tom Volk's Fungus of the Month for November 2003."
http://botit.botany.wisc.edu/toms_fungi/nov2003.html
(accessed on October 23, 2005).

**OTHER NAMES:**

See Overview

FORMULA:

- ClO_4

ELEMENTS:

Chlorine and oxygen, in combination with other ions

COMPOUND TYPE:

Inorganic salts

STATE:

Solid

MOLECULAR WEIGHT:

117.49 to 138.55 g/mol

MELTING POINT:

Ammonium perchlorate: Decomposes explosively when heated; Potassium perchlorate: 525°C (977°F); Sodium perchlorate: 480°C (896°F)

BOILING POINT:

Not applicable; all decompose at or above melting points

SOLUBILITY:

Soluble in water; ammonium perchlorate is also soluble in methyl alcohol and slightly soluble in ethyl alcohol

**K
E
Y
F
A
C
T
S**

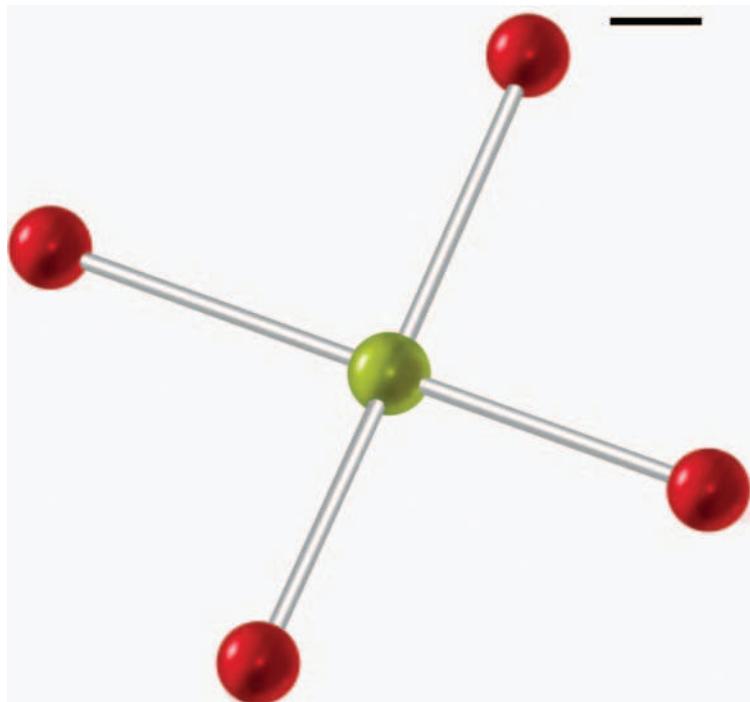
Perchlorates

OVERVIEW

The perchlorates (per-KLOR-ates) are a family of compounds consisting of salts of perchloric acid, HClO_4 . The family consists of dozens of compounds, the most important of which are ammonium perchlorate (NH_4ClO_4), potassium perchlorate (KClO_4), and sodium perchlorate (NaClO_4).

Although perchlorates have been known since the early nineteenth century, they were not produced commercially until the 1890s. Even then, they were not produced in large volumes until World War II (1939-1945), when they were made for use in explosives. Perchlorate production continued to grow in the post-war period, especially during the military build-up that accompanied the Cold War (1945-1991) between the United States and the Soviet Union. In recent years, however, concerns about the presence of perchlorates in water supplies have become increasingly widespread.

Perchlorate ion. Red atoms are oxygen. Green atom is chlorine. PUBLISHERS RESOURCE GROUP



HOW IT IS MADE

Each perchlorate is produced by a different method. Ammonium perchlorate, for example, is produced in a reaction among ammonium hydroxide (NH_4OH), hydrochloric acid (HCl), and sodium chlorate (NaClO_3).

Potassium perchlorate can be made in two ways. One method by electrolyzing potassium chlorate (KClO_3) in water. The other is by heating potassium chlorate (KClO_3). This results in a mixture of potassium perchlorate and potassium chloride (KCl), which is removed.

Sodium perchlorate is made by heating a mixture of sodium chlorate and sodium chloride. This reaction yields sodium chloride (NaCl) as well as the sodium perchlorate. The excess sodium chloride is removed from the reaction, leaving sodium perchlorate behind.

COMMON USES AND POTENTIAL HAZARDS

By far the most important uses of perchlorates are in the production of explosives and fireworks and as jet and rocket

Interesting Facts

- Small amounts of potassium perchlorate are found naturally in Chile in deposits of sodium nitrate.
- About 9 million kilograms (20 million pounds) of perchlorates are produced each year in the United States.
- In January 2005, the U.S. Environmental Protection Agency set a standard of 0.007 milligrams per kilogram of body weight as the maximum recommended dose a person should ingest of perchlorates per day.
- Between 1997 and 2004, Lockheed Martin spent \$80 million to study, clean up, and replace local water systems that had been contaminated with perchlorates used by the company in the production of military weapons and rocket fuels. The two systems were located near Riverside and Redlands, California.

fuels. These uses are based on the fact that all perchlorates are very unstable oxidizing agents. An oxidizing agent is a material that supplies oxygen to some other substances or removes electrons from that substance. When used in explosives, the perchlorates supply oxygen to the fuels in the explosive (such as sulfur or carbon), causing them to burn very rapidly, producing large volumes of gases in a very short time. When used as a rocket or jet fuel, perchlorates supply the oxygen needed to burn the fuel itself, such as kerosene or hydrogen.

Some perchlorates have other, more limited, uses. For example, potassium perchlorate was previously used to treat Graves disease, a condition in which the body produces too much thyroid hormone. It is still used to monitor the production of thyroid hormones. Potassium perchlorate is also used in emergency breathing equipment for high altitude aircraft and underwater boats. Other uses of perchlorates include:

- In nuclear reactors and electronic tubes;
- As additives in lubricating oils;
- In the tanning and finishing of leather products;

Words to Know

ELECTROLYSIS A process in which an electric current is used to bring about chemical changes.

SALT An ionic compound where the anion is derived from an acid.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

- As a fixer for fabrics and dyes;
- In electroplating;
- In the refining of aluminum metal;
- In the manufacture of rubber products;
- In the production of certain paints and enamels.

Perchlorates pose a serious risk to humans because they are unstable and have a tendency to explode spontaneously. They are also human health hazards, with harmful effects on both the brain and the thyroid. They have a tendency to prevent the uptake of iodine by the thyroid, thus interfering with the synthesis of hormones normally produced by that organ. Health officials believe that exposure to perchlorates may cause infertility in women or may have harmful effects on their newborn children. These effects include mental retardation and delays in their normal development. As a consequence, efforts are underway to locate areas where perchlorates may have entered the public water supply causing potential health problems for people living in the region.

FOR FURTHER INFORMATION

Perchlorate News.com.

<http://www.perchloratenews.com/index.html> (accessed on December 10, 2005).

"Perchlorates: New Report on Widespread Rocket Fuel Pollution in Nation's Food and Water." Organic Consumers Association. <http://www.organicconsumers.org/perchlorate.htm> (accessed on December 10, 2005).

"Potassium Perchlorate." J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/P5983.htm>
(accessed on December 10, 2005).

"Safety (MSDS) Data for Ammonium Perchlorate." Physical and Theoretical Chemistry Lab.

http://ptcl.chem.ox.ac.uk/MSDS/AM/ammonium_perchlorate.html (accessed on December 10, 2005).

Sharp, Renee, and Bill Walker. *Rocket Science: Perchlorate and the Toxic Legacy of the Cold War*. Washington, D.C.: Environmental Working Group, July 2001. Also available online at http://www.ewg.org/reports_content/rocketscience/perchlorate.pdf (accessed on December 10, 2005).

"Toxicological Profile for Perchlorates." Agency for Toxic Substances and Disease Registry.

<http://www.atsdr.cdc.gov/toxprofiles/tp162.html> (accessed on December 10, 2005).

OTHER NAMES:	Petroleum jelly; paraffin jelly; vasoliment; liquid paraffin; mineral oil; paraffin oil
FORMULA:	Not applicable
ELEMENTS:	Carbon, hydrogen
COMPOUND TYPE:	Not applicable
STATE:	Semi-solid or liquid
MOLECULE WEIGHT:	Not applicable
MELTING POINT:	Not applicable
BOILING POINT:	Not applicable
SOLUBILITY:	Insoluble in water and ethyl alcohol; soluble in benzene, chloroform, ether, carbon disulfide, and other organic solvents

KEY FACTS

Petrolatum

OVERVIEW

Petrolatum (peh-tro-LAY-tum) is a mixture, not a compound. Mixtures differ from compounds in a number of important ways. The parts making up a mixture are not chemically combined with each other, as they are in a compound. Also, mixtures have no definite composition, but consist of varying amounts of the substances from which they are formed.

Petrolatum is a complex mixture of hydrocarbons derived from the distillation of petroleum. Hydrocarbons are compounds that contain only carbon and hydrogen. The hydrocarbons that make up petrolatum belong to the methane (saturated or alkane) family of hydrocarbons with the general formula C_nH_{2n+2} . Some members of the family include methane (CH_4), ethane (C_2H_6), propane (C_3H_8), and butane (C_4H_{10}).

Petrolatum occurs in a semi-solid or liquid form. The semi-solid form is also called petroleum jelly or mineral jelly and is commercially available under a number of trade

names, including Kremoline, Pureline, Sherolatum, and VaselineTM. It ranges in color from white to yellowish to amber. It is practically odorless and tasteless. It melts over a wide range, from about 38°C to about 55°C (100°F to 131°F). The liquid form is also known as liquid paraffin, mineral oil, or white mineral oil. Such products are sold commercially under trade names such as Alboline, Drakeol, Frigol, Kremol, and Paroleine. It is a colorless, tasteless, and odorless oily liquid.

Oil was first discovered in the United States in the 1850s in western Pennsylvania. A chemist from Brooklyn, New York, Robert Augustus Chesebrough (1837–1938), visited the new wells and noticed a wax-like material sticking to the petroleum drilling rods. He learned that oil workers used the “rod wax” to heal burns on their skin. Chesebrough eventually extracted and purified the substance—petrolatum—from petroleum and began manufacturing it in 1870. He received several patents for his discovery and in 1878, he gave his product the trade name of VaselineTM. His product quickly became popular as an ointment for wounds and burns. Unlike the animal and vegetable oils then being used for that purpose, petrolatum did not spoil. By the late 1870s, VaselineTM was selling at the rate of one jar every minute in the United States. In 1880, it was added to the U.S. Pharmacopoeia, a manual that lists drugs used in medical practice.

HOW IT IS MADE

Petrolatum is a product of the fractional distillation of crude oil. Crude oil is a complex mixture of hundreds or thousands of compounds. These compounds can be separated, or distilled, from each other by heating crude oil to high temperatures. As the temperature of the crude oil rises, various groups or a “fraction” of compounds boil off. The first group of compounds includes gaseous compounds dissolved in crude oil. The next group of compounds includes compounds with low boiling points. The next group of compounds includes compounds with slightly higher boiling points. And so on. Eventually, a tar-like mass of compounds with very high boiling points is left behind in the distilling tower. This residue is heated to separate liquids from solids remaining behind. Some of these liquids and solids make up the semi-solid and liquid forms of petrolatum.

Interesting Facts

- Both solid and liquid petrolatum are available in three grades, known as USP (U.S. Pharmacopoeia), NF (National Formulary), and FCC (Food Chemicals Codex).
- A synthetic version of petrolatum is made from soybean oil as an alterna-
- tive to petroleum-based petrolatum. It is used primarily in the manufacture of cosmetics.
- Skin care products generally contain petrolatum in a concentration of about 1 to 3 percent.

COMMON USES AND POTENTIAL HAZARDS

Petrolatum has a wide variety of uses, ranging from personal care and medical applications to industrial uses. The solid form, such as Vaseline™ is used as a topical ointment for the treatment of dry, cracked skin and to reduce the risk of infection. It works as a moisturizing agent because it reduces water loss from the skin. It helps prevent infection because it creates a barrier over wounds that prevents disease-causing organisms from entering the body. Solid petrolatum is also an ingredient in many skin care and cosmetic products, such as skin lotions, body and facial cleansers, anti-perspirants, lipsticks, lip balms, sunscreens, and after-sun lotions. In hair products, it helps smooth frizzy hair by allowing hair to retain its natural moisture. The formation used in most of these products remains virtually unchanged from that developed by Robert Chesebrough in the 1800s.

Solid petrolatum is also used in industrial applications for a variety of purposes, such as:

- As a softener in the production of rubber products;
- In the food processing industry, to coat raw fruits and vegetables and to help products retain moisture;
- As a defoaming agent in the production of beet sugar and yeasts;
- For the lubrication of firearms and machine parts;

Words to Know

FRACTIONAL DISTILLATION The process of extracting compounds from petroleum by heating the petroleum and collecting the individual compounds as they boil off when their boiling points are reached.

HYDROCARBON A compound that contains hydrogen and carbon atoms.

MIXTURE A collection of two or more elements and/or compounds with no definite composition.

- In the production of modeling clays;
- In the manufacture of candles, to prevent a candle from shrinking as it cools after being burned;
- In the preparation of shoe polishes; and
- As an ingredient in rust preventatives.

The primary use of liquid petrolatum is as a laxative, a product that loosens the bowels. It also has a number of other applications, such as an additive in foods such as candies, confectionary products, and baked goods; as an ingredient in personal care products, such as baby oil creams, hair conditioning lotions, and ointments; in many different kinds of pharmaceutical preparations; in the production of industrial lubricants; as a softening agent in the manufacture of rubber, textiles, fibers, adhesives, and machine parts; as dust suppressants; and as dehydrating agents for a number of industrial processes.

FOR FURTHER INFORMATION

"Another Old Fashioned Product Vindicates Itself." *Medical Update* (October 1992): 6.

Morrison, David S. "Petrolatum: A Useful Classic." *Cosmetics and Toiletries* (January 1996): 59–68. Also available online at <http://www.penreco.com/news/events/tradearticles/petrolatumclassic.pdf> (accessed on December 22, 2005).

"Material Safety Data Sheet." Penreco.

<http://www.penreco.com/products/pdfs/petrolatum/pen00421technicalpetrolatum.pdf> (accessed on December 22, 2005).

Penreco (Petrolatum company).

<http://www.penreco.com/index.asp> (accessed on December 22, 2005).

Schramm, Daniel. "The North American USP Petrolatum Industry." *Soap & Cosmetics* (January 2002): 60-63.

See Also Petroleum

OTHER NAMES:	Crude oil; oil
FORMULA:	Not applicable
ELEMENTS:	Carbon, hydrogen, oxygen, other elements
COMPOUND TYPE:	Not applicable
STATE:	Liquid
MOLECULAR WEIGHT:	Not applicable
MELTING POINT:	Not applicable
BOILING POINT:	Not applicable
SOLUBILITY:	Not miscible with water; miscible with most organic solvents

KEY FACTS

Petroleum

OVERVIEW

Petroleum (peh-TRO-lee-yum) is a mixture, not a compound. Mixtures differ from compounds in a number of important ways. The parts that make up a mixture are not chemically combined with each other, as they are in a compound. Also, mixtures have no definite composition, but consist of varying amounts of the substances of which they are formed.

Petroleum is a very complex mixture of hydrocarbons—compounds that consist of carbon and hydrogen only. Small amounts of other organic compounds containing oxygen, sulfur, nitrogen, phosphorus, and other elements are also present. The hydrocarbons that make up petroleum include the alkanes, alkenes, alkynes, cyclic hydrocarbons, and aromatic hydrocarbons. Alkanes are hydrocarbons in which carbon and hydrogen are bonded to each other by single bonds. In alkenes, at least one double bond is present. In alkynes, carbon and hydrogen atoms are bonded by at least one triple bond. In cyclic hydrocarbons, the carbon atoms are arranged

in a ring. In aromatic hydrocarbons, the carbon atoms are joined to each other in a ring that has alternate single and double bonds. Many of the hydrocarbons in petroleum also have side chains that also consist of carbon and hydrogen atoms.

The physical properties of petroleum vary somewhat, depending on the source from which it comes and its composition. Generally, it is a thick, oily liquid that is dark yellow to brown to greenish-black in color, with a strong, unpleasant odor.

HOW IT IS MADE

Scientists believe that petroleum was formed about 300 million years ago. When microscopic plants and animals that lived in the ocean died and sank to the bottom, they were gradually covered and compressed with more layers of organic material along with sand and mud. As the mud grew thicker, it created pressure on the organic material, causing it to become increasingly warmer. The heat and pressure caused the organic matter to decay in the absence of oxygen, converting it into petroleum and natural gas.

Over time, the primitive oceans dried up. The sand and mud that had accumulated on the ocean floors changed into rock. The natural gas and liquid petroleum that had formed on the ocean floor was trapped in the rock. It flowed through cracks in the rock until it reached porous rock that acted like a sponge and soaked up the petroleum and natural gas. These fossil fuels remain trapped in the porous rock by non-porous layers of rock that act like caps or seals on the porous rocks.

Geologists have discovered a number of ways of finding these oil- and gas-soaked reservoirs. They measure changes in the Earth's magnetic field and use electronic "sniffers" that can detect hydrocarbons trapped in rock. Once oil is found, prospectors extract it by drilling into the porous rock. The pressure of gases within the porous rock pushes petroleum up to the surface, where it can be captured and stored.

Liquid petroleum has no commercial value as it comes from the earth. After it is captured from wells, therefore, it has to be refined, or separated into useful components. Refining is a process by which petroleum is heated to high temperatures in a tall cylindrical tower. The many different hydrocarbons

Interesting Facts

- The United States uses by far more petroleum than any other nation in the world. It uses about 1,000 billion liters (250 billion gallons) every year. About half that amount is imported from other countries, primarily in the Middle East.
- The average fuel economy of automobiles produced in the United States has decreased steadily since 1985.

in petroleum boil off and rise upward in the tower. The higher they rise, the more they cool off.

Each hydrocarbon eventually reaches a height at which it changes back to a liquid. Traps are inserted at various heights in the tower to catch each hydrocarbon as it changes back to a liquid. One level of traps, for example, catches hydrocarbons that condense to a liquid between about 40°C and 170°C (100°F and 340°F). This “fraction” is called the *gasoline* or *petrol* fraction of petroleum. Another level of traps catches hydrocarbons that condense between temperatures of about 170°C and 250°C (340°F and 480°F). This group of hydrocarbons is called the *kerosene* fraction. The other major fractions obtained from petroleum are the diesel and fuel oil fractions. Anything that boils above about 400°C (750°F) is referred to as the residual oil fraction. In many cases, each of the fractions obtained by this process can be further refined to separate it into even smaller fractions.

COMMON USES AND POTENTIAL HAZARDS

By far the most common use of petroleum is as gasoline for automobiles and trucks. About 90 percent of the petroleum used in the United States fuels vehicles powered by internal combustion engines. Diesel fuel is used to power trucks, buses, trains, and some automobiles. Heating oil warms buildings and powers industrial boilers. Utilities use residual fuel oil to generate electricity. Jet fuel, produced from kerosene, powers airplanes. Some airplanes use aviation

gasoline, which has a higher octane number than automobile gasoline. Octane number is a measure of a fuel's efficiency.

Kerosene is used to power lamps and heaters. Power plants and factories use petroleum coke, a solid fuel made of petroleum, as a fuel. Asphalt and tar, solid materials left over after the fractionation of petroleum, are components of paved roads. Petrolatum, another solid component of petroleum, is used as a lubricant and moisturizer. Paraffin wax, also obtained as a by-product of petroleum distillation, is an ingredient in candles, candy, polishes, and matches. Other forms of petroleum are used as lubricating oils for engines or as solvents in paints.

A second very important use of petroleum is in the manufacture of plastics and other chemicals. The number of chemical compounds obtained from petroleum and used as raw materials in chemical reactions is almost endless. It includes compounds such as methane, ethane, propane, butane, ethene (ethylene), propene (propylene), butene (butylene), benzene, methanol (methyl alcohol), ethanol (ethyl alcohol), phenol, xylene, naphthalene, and anthracene, to mention only a few.

There are many drawbacks associated with using petroleum as a fuel. One of the most important is environmental pollution. Burning petroleum products releases greenhouse gases—gases that tend to accumulate in the atmosphere and contribute to a gradual warming of the Earth's annual average temperature. Other pollutants produced when petroleum products are burned include sulfur oxides and oxides of nitrogen, compounds that contribute to the development of acid rain, and carbon monoxide and ozone, pollutants that damage plant and animal life as well as physical structures, like buildings and statues.

The production and transportation of petroleum products also poses environmental problems. Oil spills that occur during any phase of petroleum production and use may kill aquatic life and pollute an area for years. The worst oil spill in North America, for example, occurred during the wreck of the oil tanker *Exxon Valdez* in 1989, when 40,000 tons of crude oil were dumped into Alaska's Prince William Sound. The worst oil spill in the world occurred in 1978 as a result of the wreck of the oil tanker *Amoco Cadiz* off the coast of Brittany, France. More than 220,000 tons of oil were released during that accident.

Words to Know

DISTILLATION A process of separating two or more substances by boiling the mixture of which they are composed and condensing the vapors produced at different temperatures.

HYDROCARBON A compound that contains hydrogen and carbon atoms.

MISCIBLE Able to be mixed; especially applies to the mixing of one liquid with another.

MIXTURE A collection of two or more elements and/or compounds with no definite composition.

FOR FURTHER INFORMATION

American Petroleum Institute.

<http://api.ec.api.org/newsplashpage/index.cfm> (accessed on December 22, 2005).

Hyne, Norman J. *Nontechnical Guide to Petroleum Geology, Exploration, Drilling and Production*, 2nd ed. Tulsa, OK: Penn well Books, 2001.

Leffler, William L. *Petroleum Refining in Nontechnical Language*. Tulsa, OK: Pennwell Books, 2000.

“Oil and Gas Energy for the World.” Institute of Petroleum.

<http://www.energyinst.org.uk/education/oilandgas/energy.htm> (accessed on December 22, 2005).

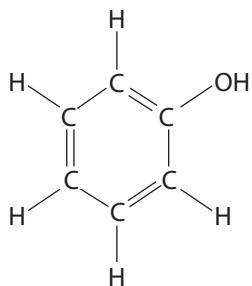
“Petroleum.” U.S. Energy Information Administration.

http://www.eia.doe.gov/oil_gas/petroleum/info_glance/petroleum.html (accessed on December 22, 2005).

“Petroleum Refining Processes.” OSHA Technical Manual.

Washington, DC: Occupational Safety and Health Administration, 1999.. Available online at
http://www.osha.gov/dts/osta/otm/otm_iv/otm_iv_2.html#1 (accessed on December 22, 2005).

See Also Methane; Petrolatum; Propane



OTHER NAMES:

Hydroxybenzene; carboxylic acid; phenyllic acid; benzophenol; phenic acid

FORMULA:

C₆H₅OH

ELEMENTS:

Carbon, hydrogen, oxygen

COMPOUND TYPE:

Aromatic alcohol (organic)

STATE:

Solid

MOLECULAR WEIGHT:

94.11 g/mol

MELTING POINT:

40.89°C (105.6°F)

BOILING POINT:

181.87°C (359.37°F)

SOLUBILITY:

Soluble in water, ethyl alcohol, ether, chloroform, acetone, benzene, and other organic solvents

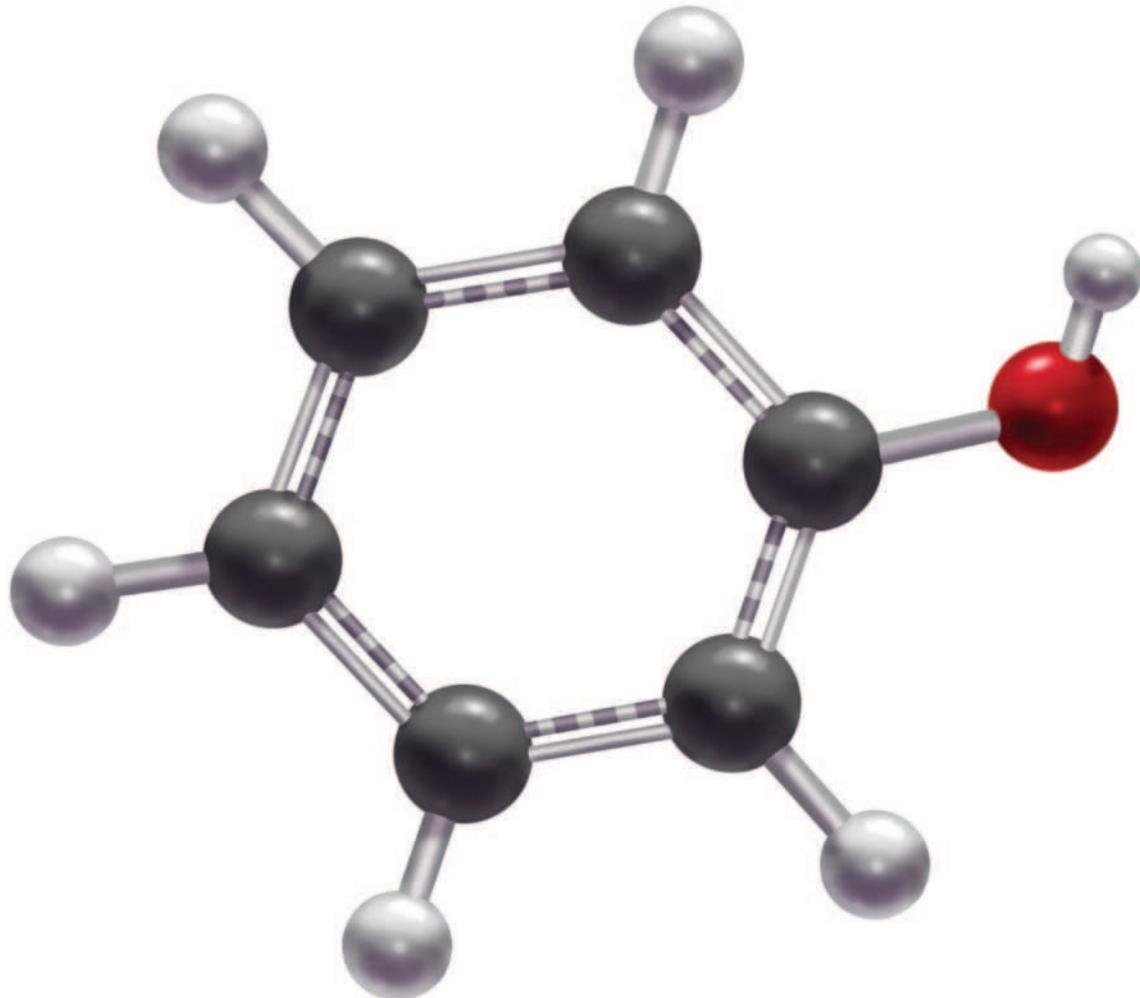
K E Y F A C T S

Phenol

OVERVIEW

Phenol (FEE-nol) is a white, crystalline solid with a characteristic odor and a sharp, burning taste. It tends to turn pink or pale red when exposed to light if not perfectly pure. Phenol has a tendency to absorb moisture from the air, changing into an aqueous solution of the compound. Such solutions have a slightly sweet flavor.

Phenol was probably first observed by German chemist Johann Rudolf Glauber (1604–1668). Glauber obtained phenol by condensing coal tar vapors and separating them into individual compounds. Coal tar is a thick black liquid left over when coal is heated in the absence of air to make coke. Phenol was largely ignored for almost 200 years until another German chemist, Friedlieb Ferdinand Runge (1795–1867), isolated phenol in 1834, also from coal tar. Runge gave the name *carboxylic acid* to his discovery, a name that is still used occasionally for the compound. In 1843, French chemist Charles Frederic Gerhardt (1816–1856) suggested the modern name of *phenol* for the compound.



Phenol. Black atoms are carbon; red atom is oxygen; white atoms are hydrogen.
Bonds in the benzene ring are represented by the double striped sticks. White sticks show single bonds. PUBLISHERS RESOURCE GROUP

The term *phenol* is also used to describe a class of aromatic organic compounds. Such compounds have a six-carbon ring structure like that of benzene (C_6H_6) to which is attached one or more hydroxyl (-OH) groups. Benzene is the simplest member of this group. Other members include the cresols ($CH_3C_6H_4OH$), xylenols ($(CH_3)_2C_6H_3OH$), and the resorcinols ($C_6H_4(OH)_2$). The production of phenol amounts to about 6 million metric tons (about 6.5 million short tons) annually in the United States. The largest single use of phenol

Interesting Facts

- The only source of phenol until World War I (1914–1918) was coal tar.
- German chemist Friedrich Raschig (1863–1928) developed the first synthetic method for making phenol in 1915. The process involves the hydrolysis of chlorobenzene (C_6H_5Cl).
- Demand for phenol increased significantly in 1909 when German chemist Leo Hendrik Baekeland (1863–1944) announced the discovery of a new synthetic material, Bakelite, made from phenol and formaldehyde (CH_2O). Bakelite was the first thermosetting plastic ever discovered. A thermosetting plastic is one that, once shaped into some form, cannot be remelted again.
- The use of carbolic acid as an antiseptic was first suggested by Sir Joseph Lister (1827–1912) in 1865. Lister found that sterilizing medical instruments and treating wounds with carbolic acid dramatically reduced the number of deaths resulting from post-operative infections.
- The effectiveness of a substance in preventing bacterial infection is still expressed by a measure known as the *phenol coefficient*, the effectiveness of the material compared to that of a comparable amount of phenol.

is in the production of bisphenol A ($(CH_3)_2C(C_6H_4OH)_2$) and a variety of plastics, primarily formaldehyde resins.

HOW IT IS MADE

In the early twenty-first century, virtually all of the phenol made is produced from cumene ($C_6H_5CH(CH_3)_2$). Cumene is first oxidized to produce cumene hydroperoxide: $C_6H_5CH(CH_3)_2 + O_2 \rightarrow C_6H_5(COOH)(CH_3)_2$. The cumene hydroperoxide is then treated with concentrated sulfuric acid (H_2SO_4) to obtain a mixture of acetone (CH_3COCH_3) and phenol: $C_6H_5(COOH)(CH_3)_2 + H_2SO_4 \rightarrow CH_3COCH_3 + C_6H_5OH$.

COMMON USES AND POTENTIAL HAZARDS

The primary use for phenol is as an intermediary chemical, a compound used in the synthesis of other chemicals. About 40 percent of all the phenol produced in the United States is used to make bisphenol A, while a similar amount is used in the production of a variety of polymers, such as phenol-formaldehyde plastics and nylon-6. The third largest application of phenol is in the manufacture of a host of other chemicals, xylene and aniline being the most important.

Phenol is no longer widely used as an antiseptic, partly because more efficient substances have been developed and partly because phenol may cause irritation and burning of the skin after prolonged use. The compound is still used in low concentrations in a number of health and medical applications, however, as in antipruritics (substances that reduce or prevent itching), cauterizing agents (substances for the burning of tissue by heat or chemicals), topical anesthetics (anesthetics used on the skin), chemexfoliants (substances that remove skin), throat sprays and lozenges (such as Chloraseptic®, Ambesol®, and Cepastat®), and in skin ointments (such as PRID salve® and CamphoPhenique® lotion). Phenol is also used in combination with slaked lime (calcium hydroxide; $\text{Ca}(\text{OH})_2$) and other materials as a disinfectant for toilets, stables, cesspools, floors, and drains.

Phenol is a highly corrosive material that can cause serious burns to the skin, eyes, and respiratory system. It is toxic if ingested. The compound can enter the body in a variety of ways, such as absorption through the skin, inhalation of vapors, and ingestion of the solid compound or its solutions. Symptoms of phenol poisoning include nausea, vomiting, headache, respiratory failure, muscular weakness, severe depression, collapse, coma, and death. Skin exposure may cause redness, blisters, and/or minor to severe chemical burns. Ingestion of phenol may cause damage to the central nervous system, lungs, kidneys, liver, pancreas, and spleen. The concentration of phenol in most consumer and industrial products is low enough not to cause health problems for users. However, prolonged exposure to such products or their overuse may result in serious health risks. Handling of the pure compound or concentrated solutions of phenol also involves health hazards.

Words to Know

AQUEOUS SOLUTION A solution that consists of some material dissolved in water.

HYDROLYSIS The process by which a compound reacts with water to form two new compounds.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

FOR FURTHER INFORMATION

“Occupational Safety and Health Guideline for Phenol.” Occupational Safety and Health Administration.

<http://www.osha.gov/SLTC/healthguidelines/phenol/recognition.html> (accessed on December 29, 2005).

“Phenol.” Greener Industry.

http://www.uyseg.org/greener_industry/pages/phenol/1PhenolAnnualProd.htm (accessed on December 29, 2005).

“Phenol.” International Programme on Chemical Safety.

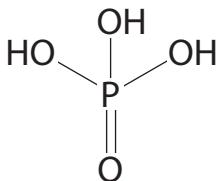
<http://www.inchem.org/documents/pims/chemical/pim412.htm> (accessed on December 29, 2005).

“ToxFAQsTM for Phenol.” Agency for Toxic Substances and Disease Registry.

<http://www.atsdr.cdc.gov/tfacts115.html> (accessed on December 29, 2005).

Tyman, J. H. P. *Synthetic and Natural Phenols*. Amsterdam: Elsevier, 1996.

See Also Cumene; Polyamide 6,6



OTHER NAMES:	Orthophosphoric acid
FORMULA:	H_3PO_4
ELEMENTS:	Hydrogen, phosphorus, oxygen
COMPOUND TYPE:	Inorganic acid
STATE:	Solid. See Overview
MOLECULAR WEIGHT:	98.00 g/mol
MELTING POINT:	42.4°C (108°F)
BOILING POINT:	407°C (765°F)
SOLUBILITY:	Very soluble in water and ethyl alcohol

K E Y F A C T S

Phosphoric Acid

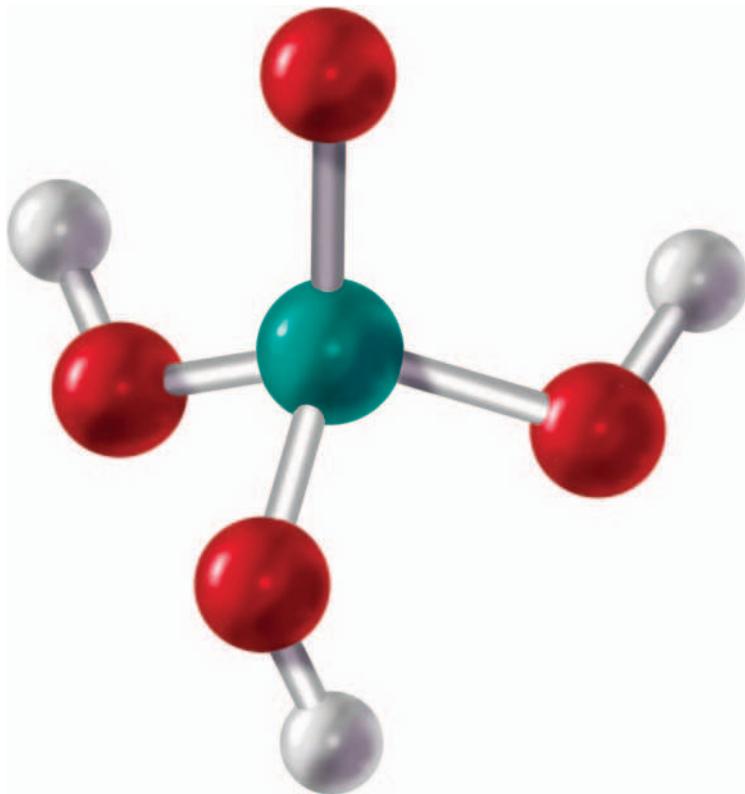
OVERVIEW

Phosphoric acid (fos-FOR-ik AS-id) melts at a temperature just above room temperature (about 20°C; 68°F), so would be expected to occur as a solid under those conditions. As a solid, the acid is a white crystalline material with a strong tendency to absorb moisture from the air. In fact, phosphoric acid may also occur as a supercooled liquid at room temperature. A supercooled liquid is one that remains in a liquid state at temperatures below its freezing point, at which temperature it would be expected to be a solid. As a liquid, phosphoric acid is a colorless, odorless, syrupy liquid whose character is sometimes described as sparkling.

Phosphoric acid was discovered independently as a component of bone ash in 1770 by two Swedish chemists, Johann Gottlieb Gahn (1745–1818) and Karl Wilhelm Scheele (1742–1786). Four years later, Scheele discovered that the acid could be made by adding nitric acid to phosphorus.

Phosphoric acid is the ninth highest volume chemical produced in the United States. In 2004, the U.S. chemical

Phosphoric acid. Red atoms are oxygen; white atoms are hydrogen; turquoise atom is phosphorus. Gray stick shows a double bond. PUBLISHERS RESOURCE GROUP



industry made about 5.2 million kilograms (11.5 million pounds) of phosphoric acid. About 90 percent of that amount went to the manufacture of fertilizers.

HOW IT IS MADE

The most economical method for making phosphoric acid is by treating phosphate rock with sulfuric acid (H_2SO_4). Phosphate rock is naturally-occurring rock with large amounts of calcium phosphate ($Ca_3(PO_4)_2$). The product of this reaction is generally not very pure, but sufficiently pure for use in the production of fertilizers. Higher quality phosphoric acid can be made by burning phosphorus or phosphate rock in an electric furnace, converting either to gaseous phosphoric oxide (diphosphorus pentoxide; P_2O_5). The oxide is then dissolved in water to form the acid.

Interesting Facts

- Phosphoric acid is commercially available in a number of technical grades, ranging from agricultural (relatively low purity) to technical (from 50 to 100 percent purity) to FCC (Food Chemicals Codes) quality of at least 75 percent phosphoric acid.

COMMON USES AND POTENTIAL HAZARDS

By far the most important use of phosphoric acid is in the production of fertilizers. At least four major types of fertilizers are made from phosphoric acid: diammonium phosphate ($(\text{NH}_4)_2\text{HPO}_4$; DAP), monoammonium phosphate ($\text{NH}_4\text{H}_2\text{PO}_4$; MAP), granulated triple superphosphate (GTSP), and superphosphoric acid, the only liquid among the group. An additional 5 percent of the phosphoric acid produced is used as an animal feed supplement.

The remaining 5 percent of phosphoric acid produced is used in a very wide range of commercial, industrial, and household products, including:

- For pickling, cleaning and treating metallic surfaces, especially in the steel industry;
- In the synthesis of inorganic chemical compounds;
- As a catalyst in the manufacture of ethanol (ethyl alcohol), ethylene, and other organic compounds;
- As a food additive in a number of products, such as colas, beers, jams, and cheeses, where it adds a touch of tartness to the product;
- In the dentistry profession, where it is used to etch and clean teeth;
- In a number of consumer products, such as soaps, detergents, and toothpastes;
- As a refining and clarifying agent in the production of sugar;
- In the dyeing of cotton;

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

PHOSPHATE A compound that is manufactured from phosphoric acid

SUPERCOOLED Refers to a substance that remains in a liquid state at temperatures below its freezing point.

- As a binder for cement;
- In the manufacture of waxes and polishes; and
- In water and sewage treatment plants.

Phosphates (compounds made from phosphoric acid) were once used widely as “builders” in detergents. A builder is a compound that increases the cleaning efficiency of the detergent. The problem is that phosphates that escape into the natural environment can result in some undesirable changes in fresh water systems. Algae living in these systems use phosphate to grow and multiply, resulting in the conversion of fresh water lakes and ponds into swamps and bogs, and, eventually, into dry fields, a process known as eutrophication. Because of this effect, the use of phosphates in detergents has been banned in most parts of the United States.

Phosphoric acid is an extremely hazardous and toxic compound. In small amounts, it causes irritation of the skin, eyes, and respiratory system. If ingested, it can cause serious damage to the digestive system, resulting in nausea, vomiting, abdominal pain, difficulty in breathing, shock, and occasionally death by asphyxiation (suffocation). The most serious health hazards posed by phosphoric acid are of concern primarily to people who work with the product. The amount of phosphoric acid present in most domestic and household products is very small and poses little risk to users of those products.

FOR FURTHER INFORMATION

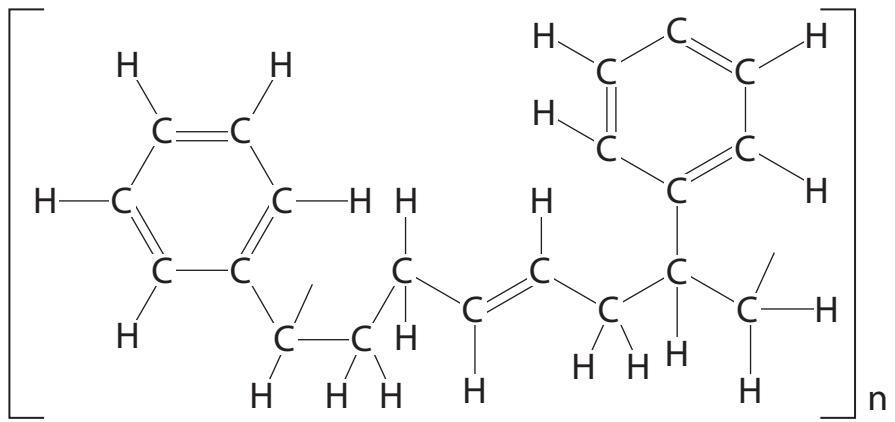
“Chemical of the Week: Phosphoric Acid.” Science Is Fun.
<http://scifun.chem.wisc.edu/CHEMWEEK/H3PO4/H3PO4.html>
(accessed on October 24, 2005).

"Phosphate Primer." Florida Institute of Phosphate Research.
<http://www1.fipr.state.fl.us/PhosphatePrimer> (accessed on October 24, 2005).

"Phosphoric Acid." DC Chemical Co., Ltd.
http://www.dcchem.co.kr/english/product/p_basic/p_basico4.htm (accessed on October 24, 2005).

"Phosphoric Acid Fuel Cells." Smithsonian Institution.
<http://americanhistory.si.edu/fuelcells/phos/pafcmain.htm> (accessed on October 24, 2005).

See Also Nitric Acid; Sulfuric Acid



OTHER NAMES:
SBS

FORMULA:
 $-\text{[CH}_2\text{CHC}_6\text{H}_5\text{]}_n-$
 $-\text{[CH}_2\text{CH=CHCH}_2\text{]}_n-$
 $[\text{CH}_2\text{CHC}_6\text{H}_5\text{]}_n$

ELEMENTS:
Carbon, hydrogen

COMPOUND TYPE:
Organic polymer

STATE:
Solid

MOLECULAR WEIGHT:
Varies

MELTING POINT:
160°C to 200°C
(320°F to 400°F)

BOILING POINT:
Not applicable

SOLUBILITY:
Insoluble in water

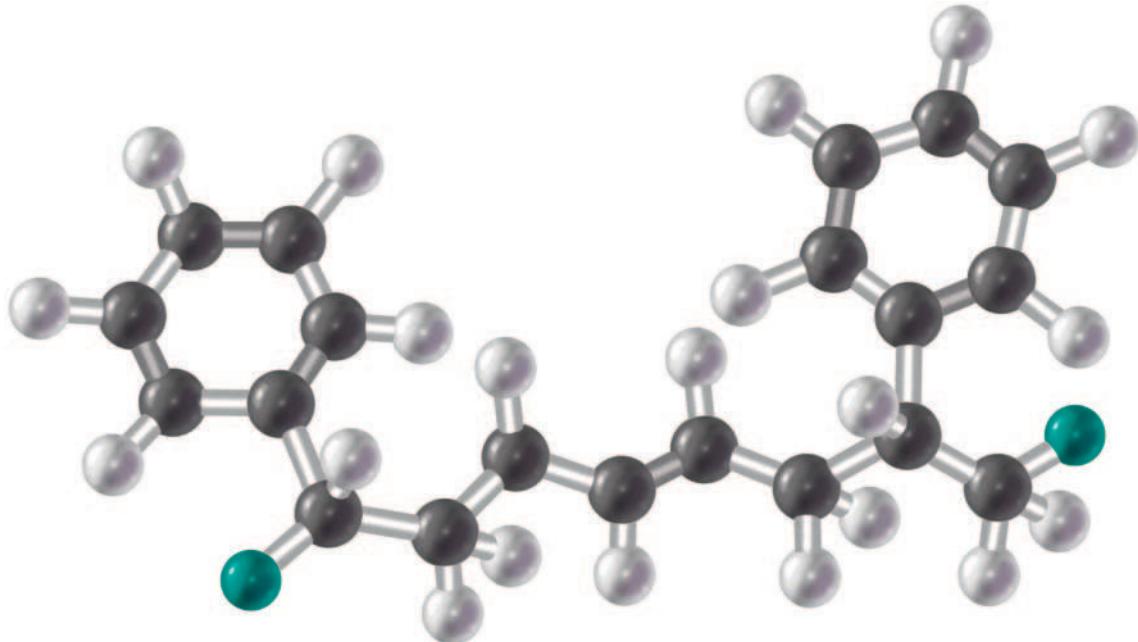
KEY FACTS

Poly(Styrene-Butadiene-Styrene)

OVERVIEW

Poly(styrene-butadiene-styrene) (pol-ee-STYE-reen-byoo-tah-DYE-een-STYE-reen) is a thermoplastic block copolymer of styrene and butadiene. The compound is often called simply SBS or SBS rubber. A thermoplastic polymer is one that can be converted back and forth between liquid and solid states by alternate heating and cooling. A copolymer is a polymer made from two monomers, in this case, styrene ($\text{C}_6\text{H}_5\text{CH=CH}_2$) and 1,3-butadiene ($\text{CH}_2=\text{CHCH=CH}_2$). The term *block copolymer* means that one section of the polymer chain consists of polystyrene ($-\text{[CH}_2\text{CHC}_6\text{H}_5\text{]}_n$) to which is connected another section consisting of polybutadiene ($-\text{[CH}_2\text{CH=CHCH}_2\text{]}_n$), which, in turn, is connected to another section of polystyrene ($-\text{[CH}_2\text{CHC}_6\text{H}_5\text{]}_n$), and so on.

A copolymer like SBS has properties of both polymers of which it is composed. In the case of SBS, the polystyrene segments give the product strength and durability, while the polybutadiene segments provide flexibility. The substance acts like natural rubber at room temperature, but becomes



Poly (styrene butadiene styrene). White atoms are hydrogen; black atoms are carbon; and turquoise atoms show where this molecule joins to other ones to form chains.

PUBLISHERS RESOURCE
GROUP

soft and plastic when heated. The latter property means that products made of SBS can be formed into a variety of shapes.

SBS is resistant to abrasion and does not readily break down when exposed to heat, light, and chemicals. It may dissolve or break down when exposed to fats and oils and various types of hydrocarbon compounds and mixtures. It maintains its structure and performance well over a wide temperature range from -60°C to 150°C (-75°F to 300°F).

SBS was first developed in the early 1930s by two German chemists, Walter Bock and Eduard Tschunkur. Their research was part of Germany's Four Year Plan for self-sufficiency. Under that plan, the nation worked toward eliminating, so far as possible, the import of essential materials that might be needed in case of a war. Germans already had one type of synthetic rubber, known as *Buna*, but it had a number of disadvantages. SBS was far superior to *Buna*, and was soon being produced in very large amounts in German factories.

HOW IT IS MADE

Molecules of both styrene and 1,3-butadiene contain double bonds. Any compound with double bonds has the ability

to form polymers. Polymerization occurs when the double bond in one monomer molecule (such as styrene) breaks apart. A hydrogen atom from a second molecule of the monomer then adds on to one end of the broken double bond. The rest of the second molecule adds to the other end of the broken double bond. A “double-molecule,” consisting of two monomers joined to each other, forms. The “double-molecule” also contains a double bond. So the process can be repeated to form a “triple-molecule” consisting of three monomer molecules. The process is repeated hundreds or thousands of times producing a long chain of monomers.

In the production of a block copolymer, one extra step is added. First, a long chain of styrene monomers is produced. Then a long chain of butadiene monomers is made. Next, the two chains are joined to each other. Finally, additional chains of polystyrene and polybutadiene are added, making a very long chain consisting of alternating blocks of polystyrene and polybutadiene.

This method is used for the production of many different kinds of polymers. The most difficult problems may be (1) how to get the first few double bonds to break apart, and (2) how to stop the polymerization reaction at just the right point. The method used to make most SBS today involves the use of a butyl lithium (C_4H_9Li) catalyst, which is very effective in getting the reaction started. The reaction is terminated at some given point by the addition of dichlorodimethylsilane ($SiCl_2(CH_3)_2$). The dichlorodimethylsilane reacts with the last monomer at the end of the SBS chain, blocking the addition of any additional styrene or butadiene monomers.

COMMON USES AND POTENTIAL HAZARDS

The process by which SBS is made can be modified to produce products with somewhat different physical and chemical properties. For example, some forms of SBS are especially strong, making them suitable for the manufacture of tires, shoe soles, conveyer belts, and the tracks on caterpillar trucks. Other types of SBS are engineered to be more flexible, for use as rubber tubing, flexible toys, sporting goods, and refrigerator gaskets. SBS products can also be made in a variety of colors and shapes for use as seals, rubber mats, floor coverings, tire treads, and shoe components.

Words to Know

COPOLYMER A polymer made with two different monomers.

MONOMER A small molecular unit that joins with others to form a polymer.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

THERMOPLASTIC A material that becomes soft and moldable when heated, then hardens when it is cooled.

SBS production fluctuates or changes according to a number of factors, including market demand, the price of petroleum, and the price of natural rubber. For example, when natural rubber is readily available and inexpensive, the demand for synthetic types of rubber, such as SBS, decreases. Also, when the price of petroleum increases, SBS becomes more expensive to make and production decreases.

FOR FURTHER INFORMATION

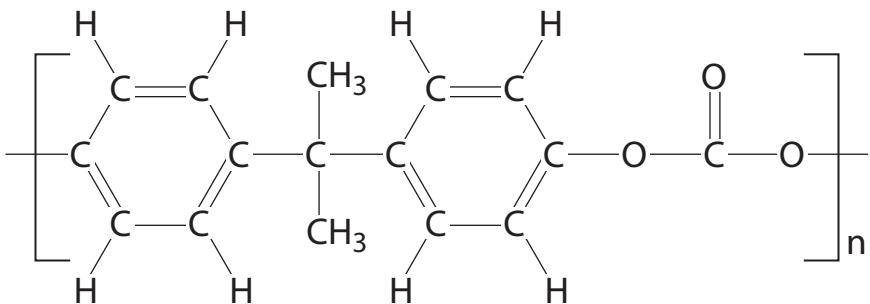
Dick, John S., and R. A. Annicelli, eds. *Rubber Technology: Compounding and Testing for Performance*. Cincinnati, OH: Hanser Gardner Publications, 2001.

Johnson, Peter S. *Rubber Processing: An Introduction*. Cincinnati, OH: Hanser Gardner Publications, 2001.

“SBS Rubber.” University of Mississippi, Polymer Science Learning Center.

<http://www.psfc.ws/mactest.sbs.htm> (accessed on October 26, 2005).

See Also 1,3-Butadiene; Polystyrene; Styrene



FORMULA:
Varies

ELEMENTS:
Carbon, hydrogen,
oxygen

COMPOUND TYPE:
Organic polymer

STATE:
Solid

MOLECULAR WEIGHT:
Very large; varies

MELTING POINT:
Varies

BOILING POINT:
Not applicable;
decomposes above
melting point

SOLUBILITY:
Virtually insoluble in
water

KEY FACTS

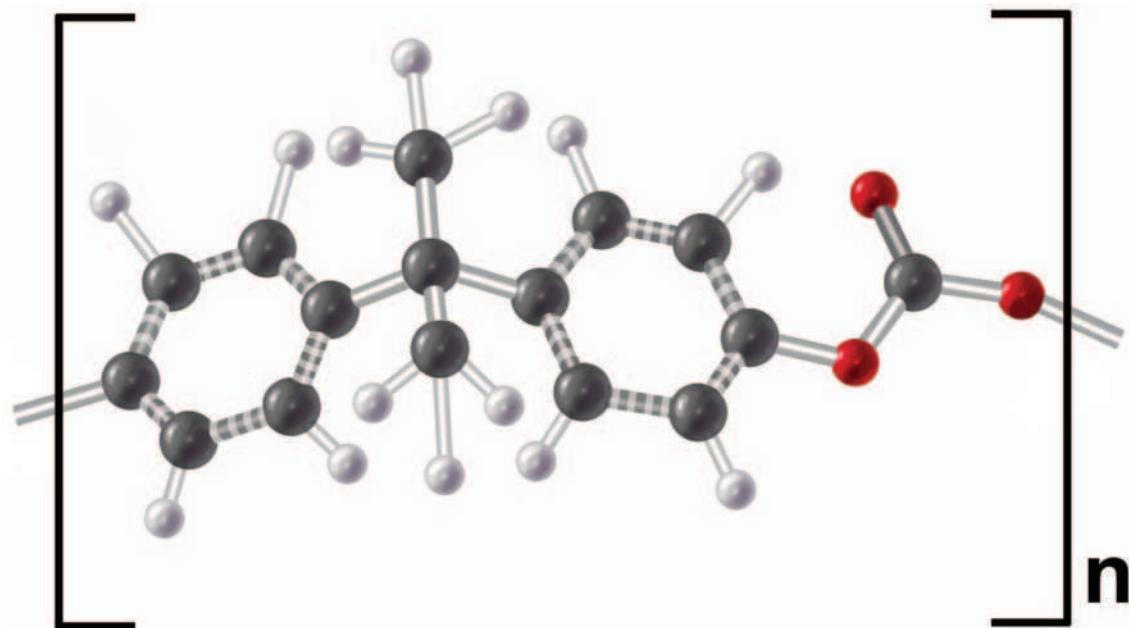
Polycarbonates

OVERVIEW

Polycarbonate (pol-ee-KAR-bun-ate) is a term used both for a specific compound and for a class of compounds with similar chemical structures. The members of this family are made by reacting phosgene (COCl_2) with any compound having two phenol structures. Phenol is hydroxybenzene, $\text{C}_6\text{H}_5\text{OH}$. The most common polycarbonate is made in the reaction between phosgene and bisphenol A ($\text{C}_6\text{H}_5\text{OHC(CH}_3)_2\text{C}_6\text{H}_5\text{OH}$). Polycarbonates are sold under a number of trade names, including Cyrolon®, Lexan®, Markrolon®, Merlon®, Tuffak®, and Zelux®.

Polycarbonates are strong, lightweight plastics that are resistant to heat, light, chemicals, and physical shock. They are used in a number of commercial and industrial products ranging from consumer electronics to sporting goods to storage containers.

The original polycarbonate was discovered in 1953 by the German chemist Hermann Schnell (1916–1999), an employee at the Bayer AG chemical company. The product was discovered



Polycarbonate. Red atoms are oxygen; black atoms are carbon; white atoms are hydrogen. Gray sticks indicate double bonds. Striped sticks show benzene rings. PUBLISHERS RESOURCE GROUP

almost simultaneously in the United States by chemist Daniel W. Fox (1923–1989) at the General Electric Company. Commercial use of the product began in the late 1950s, with the first large-scale production plant beginning operation in 1960. At first, polycarbonate was used primarily for electrical devices such as fuse boxes. In 1982, compact discs (CDs) were introduced and were made almost entirely of polycarbonate. Fifteen years later, polycarbonate digital videodiscs (DVDs) entered the market. In the mid-1980s, polycarbonate bottles began to replace the more cumbersome and breakable glass bottles. Today, polycarbonates are used for dozens of commercial applications.

HOW IT IS MADE

Polycarbonates are made in the reaction between phosgene and any compound containing two phenol groups, such as bisphenol A. One chlorine atom from the phosgene combines with a hydrogen atom from the hydroxyl group (-OH) in each phenol to make hydrogen chloride (HCl). As the hydrogen chloride is removed from the reaction, the phosgene remnant links with the phenol remnant. As the reaction proceeds, the phosgene-phenol grouping grows longer and

Interesting Facts

One of polycarbonate plastics most interesting uses is in the manufacture of bulletproof windows.

longer, eventually forming a large, straight-chain polymer of polycarbonate.

Other types of polycarbonates have also been made using a very different approach from that involving bisphenol A and related compounds. For example, the reaction between phosgene and allyl alcohol ($\text{CH}_2=\text{CHCH}_2\text{OH}$) produces a monomer with carbon-carbon double bonds at both ends of the molecule that can be used for polymerization. Interestingly enough, the polycarbonate produced by this process has very different physical properties from the traditional bisphenol A polymer. The allyl polymer is a clear, transparent, flexible plastic whose primary use is in the production of eyeglass lenses.

COMMON USES AND POTENTIAL HAZARDS

Polycarbonates are strong, heat resistant, and lightweight, making them ideal for applications in construction, electronics, automobiles, and appliances. Many polycarbonates are substituted for glass in safety and athletic goggles, building components, and car instrument panels because they are transparent, yet shatterproof. They are also more resistant to ultraviolet radiation than is glass. Some of the specific products made with polycarbonates include:

- Consumer electronic devices, such as cell phones, computers, pagers, and fax machines;
- Data storage devices, such as CDs and DVDs;
- Automobile parts, including tail lights, turn signals, fog lights, and headlamps;
- Appliances such as refrigerators, food mixers, hair dryers, and electric shavers;
- Safety devices, including helmets, goggles, and bullet-proof windows;

Words to Know

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

- Healthcare devices, such as incubators, kidney dialysis machines, and eyeglasses.

The U.S. Food and Drug Administration has accepted polycarbonates as safe for use with foods. There is little evidence that the compounds have any harmful health effects on humans. Minimal evidence exists to suggest that bisphenol A may have some harmful effects on experimental animals. But there is no evidence that similar effects occur in humans, and the amount of free bisphenol A in polycarbonates is so small as to be negligible.

FOR FURTHER INFORMATION

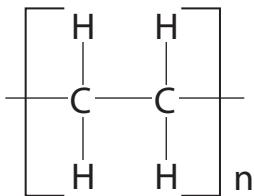
Lazear, N. R. "Polycarbonate: High Performance Resin." *Advanced Materials & Processes* (February 1995): 43–45.

"Polycarbonate." Association of Plastics Manufacturers in Europe. <http://www.plasticseurope.org/content/default.asp?PageID=43> (accessed on October 24, 2005).

"Polycarbonate (Makrolon®, Lexan®, Zelux®)." San Diego Plastics. <http://www.sdplastics.com/polycarb.html> (accessed on October 24, 2005).

"Polycarbonate Plastics." Bisphenol A. <http://www.bisphenol a.org/human/polyplastics.html> (accessed on October 24, 2005).

"Polycarbonates." Polymer Science Learning Center. <http://www.pslc.ws/mactest/pc.htm> (accessed on October 24, 2005).



OTHER NAMES:	Ethene homopolymer
FORMULA:	$[-\text{CH}_2-\text{CH}_2-]_n$
ELEMENTS:	Carbon, hydrogen
COMPOUND TYPE:	Organic polymer
STATE:	Solid
MOLECULAR WEIGHT:	1,500 to 100,000 g/mol
MELTING POINT:	Varies: about 85°C -110°C (185°F-230°F)
BOILING POINT:	Not applicable; decomposes above melting point
SOLUBILITY:	Insoluble in water and most organic solvents; soluble in hydrocarbons and halogenated hydrocarbons

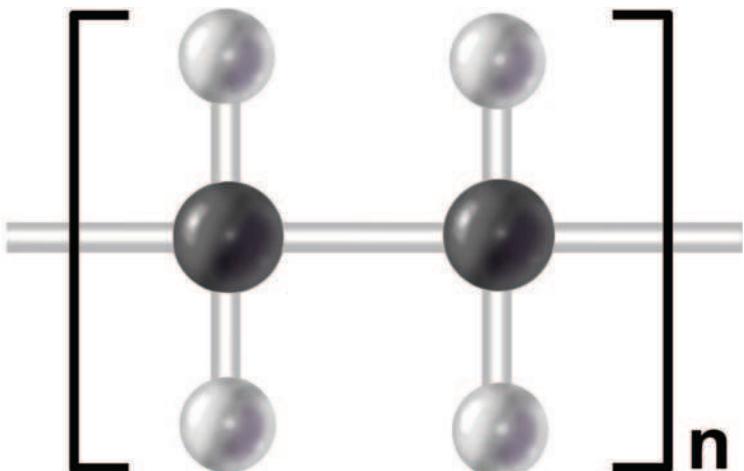
KEY FACTS

Polyethylene

OVERVIEW

Polyethylene (pol-ee-ETH-uh-leen) is a thermosetting white solid resistant to high temperatures, most inorganic and organic chemicals, and physical impact. It is also an electrical non-conductor. A thermosetting polymer is one that, once it is melted and formed, can not be re-melted. Polyethylene is available in a variety of forms, the most common of which are high-density (HD or HDPE), low density (LD or LDPE), linear low density (LLD or LLDPE) and cross-linked (CLPE). These forms of the compound differ with respect to the structure of the polyethylene chains and their relationship to each other. For example, if all of the polyethylene chains are straight chains without branches, they can pack together tightly forming a high density product. By contrast, low density polyethylene consists of shorter chains with many side branches on them. The side branches prevent adjacent polymer chains from getting too close to each other. In cross-linked polyethylene, adjacent polymer chains actually form chemical bonds with each other, holding them in a regular, almost crystalline pattern.

Polyethylene. White atoms are hydrogen and black atoms are carbon. PUBLISHERS RESOURCE GROUP



Polyethylene was first prepared accidentally in 1889 by the German chemist Hans von Pechmann (1850-1904). Von Pechmann was heating diazomethane ($\text{H}_2\text{C}=\text{N}=\text{N}$) when he observed the formation of a white, waxy solid. When his colleagues identified the substance as containing repeated methylene (-CH₂) units, they called the material polymethylene. Von Pechmann did not pursue his discovery, nor did any of his colleagues until the 1930s. Then, in 1933, two chemists at the English firm of Imperial Chemicals Industries (ICI), Reginald Gibson and Eric Fawcett, accidentally re-discovered polyethylene. While attempting to pressurize a mixture of ethylene and benzaldehyde, Gibson and Fawcett observed the formation of a white waxy solid within the pressure vessel. At first, they were unable to account for this reaction. Eventually, however, they discovered that a tiny hole in the pressure vessel had allowed oxygen to seep into the tank, catalyzing the conversion of ethylene to polyethylene. Two years later, two other ICI chemists, J. C. Swallow and M. W. Perrin duplicated the Gibson-Fawcett experiment and devised an efficient commercial method for making the polymer.

Beginning in the 1950s, researchers in a number of countries began to explore the use of catalysts to increase the efficiency of reactions by which polyethylene is made. For example, Robert Banks and John Hogan at Phillips Petroleum invented a procedure using chromium trioxide (Cr_2O_3) for the preparation of high-density polyethylene and another form of the product known as crystalline polyethylene. In 1953, the German chemist Karl Ziegler (1898-1973)

Interesting Facts

Phillips Petroleum had difficulty maintaining quality control in the early years of making polyethylene. As a result, it produced large quantities of the product that could not be sold for commercial, industrial, or household use. The company faced financial ruin. Fortunately, a new toy came into existence in the mid 1950s, the hula hoop. A hula hoop is simply a ring of plastic that

could be made with low grade polyethylene. The first company to market hula hoops, called Wham-O, bought up much of Phillips' defective polyethylene stock for its hula hoops. The company sold more than twenty million hula hoops in its first six months of existence at a cost of \$1.98 each. Wham-O was a booming success, and it saved Phillips from financial ruin.

used titanium halides and organoaluminum compounds to make polyethylene under even lower temperatures and pressures. And in 1976, the German chemists Walter Kaminsky (1941-) and Hansjörg Sinn (1929-) invented a third method of production using metallocenes (organic compounds that contain a metal) as a catalyst.

HOW IT IS MADE

Polyethylene is made by polymerizing ethylene (ethene; $\text{CH}_2=\text{CH}_2$). Polymerization occurs when the double bond in ethylene breaks, allowing one molecule of ethylene to combine with a second molecule of ethylene: $\text{CH}_2=\text{CH}_2 + \text{CH}_2=\text{CH}_2 \rightarrow \text{CH}_3\text{CH}_2\text{CH}=\text{CH}_2$. The product of that reaction also contains a double bond, allowing the reaction to be repeated: $\text{CH}_3\text{CH}_2\text{CH}=\text{CH}_2 + \text{CH}_2=\text{CH}_2 \rightarrow \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}=\text{CH}_2$. Once again, the final product contains a double bond, and the reaction can be repeated again and again and again.

Polymerization occurs when some outside agent provides the energy to break the double bond in ethylene to get the reaction started. Heat, light, ultraviolet radiation, a beam of electrons, and gamma rays have all been used to initiate polymerization. Polymerization occurs at lower energy levels, and it has been the search for catalysts to achieve this objective

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

that have led to the processes developed by Banks and Hogan, Ziegler, Kaminsky and Sinn, and other researchers.

COMMON USES AND POTENTIAL HAZARDS

An estimated 55 million metric tons (60 million short tons) of polyethylene are produced worldwide each year. In the United States, an estimated 7 million metric tons (8 million short tons) of high-density polyethylene will be produced in 2006 and an estimated 2.7 million metric tons (3.0 million short tons) of low-density polyethylene will be made. The greatest fraction of HDPE is used in the manufacture of molded products, film and sheeting, pipes and tubing, fibers, and gasoline and oil containers. The most important application of LDPE is in the manufacture of packaging films for foods as well as coatings that are sprayed or otherwise applied to all kinds of surfaces. It is also used widely to make liners for drums and other shipping containers, wire and cable coating, trash bags, squeeze bottles, inexpensive dinnerware, drop cloths, swimming pool covers, toys, and electric insulation.

No health hazards from polyethylene in any form have yet been identified.

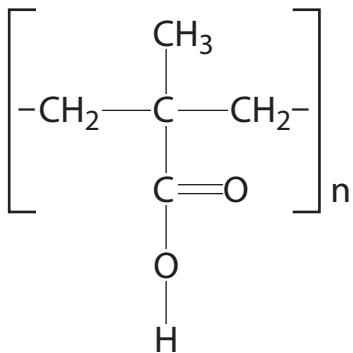
FOR FURTHER INFORMATION

Meikle, Jeffrey L. *American Plastic: A Cultural History*.
http://www.iilo.org/public/english/protection/safework/cis/products/icsc/dtasht_icsc14/icsc1488.htm (accessed on October 24, 2005).

“Polyethylene.” In *World of Invention*. 2nd ed. Edited by Kimberley A. McGrath and Bridget Travers. Detroit, MI: Gale, 1999.

“Polyethylene Specifications.” Boedeker Plastics. http://www.boedeker.com/polye_p.htm (accessed on October 24, 2005).

See Also Polypropylene



OTHER NAMES:
Acrylic, PMMA

FORMULA:
 $-\text{[}-\text{CH}_2\text{C}(\text{CH}_3)\text{(COOH)}\text{CH}_2\text{-]}_n$

ELEMENTS:
Carbon, hydrogen,
oxygen

COMPOUND TYPE:
Organic polymer

STATE:
Solid

MOLECULAR WEIGHT:
Varies: 250,000 to
over 1,000,000 g/mol

MELTING POINT:
Varies: usually above
100°C (200°F)

BOILING POINT:
Not applicable

SOLUBILITY:
Insoluble in water;
best solvents are
mixtures of two or
more organic
solvents, aromatic
hydrocarbons,
halogenated hydro-
carbons, and
tetrahydrofuran

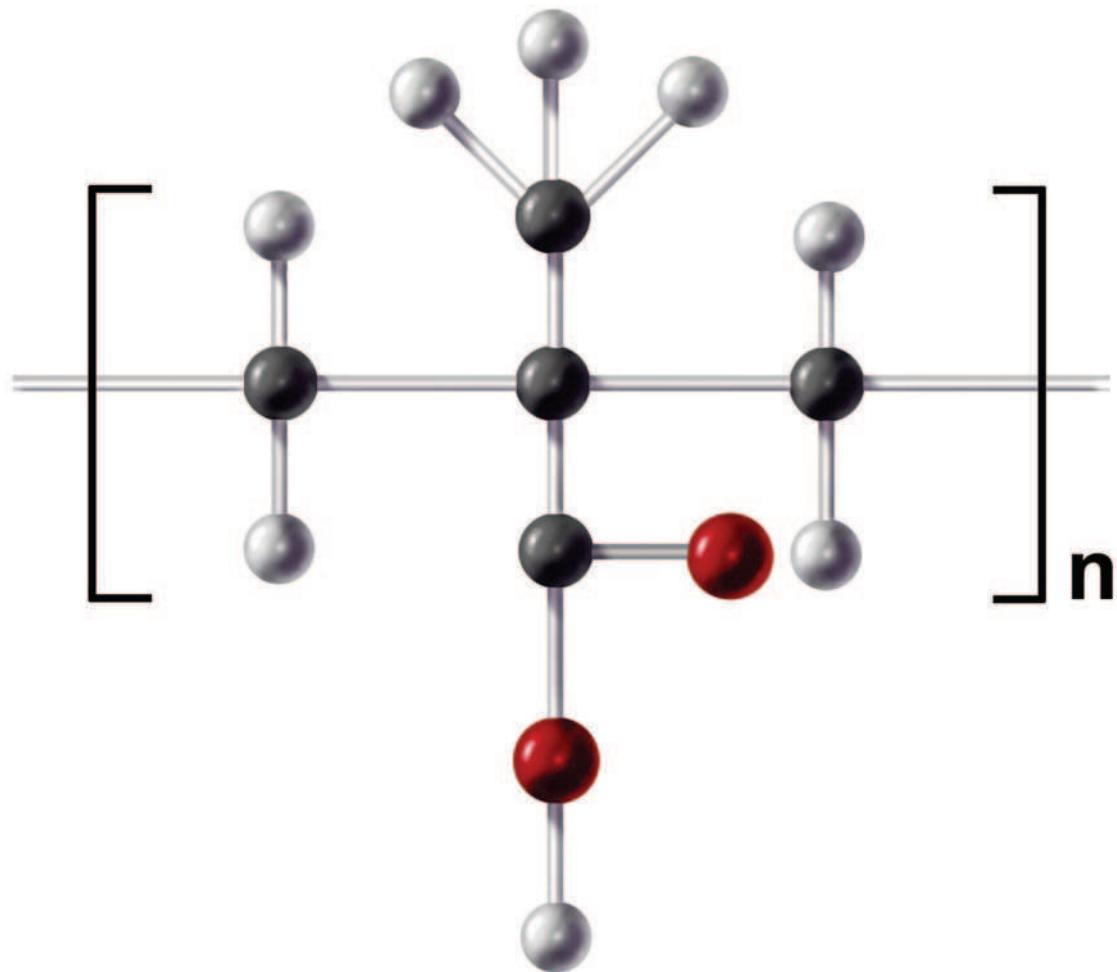
KEY FACTS

OVERVIEW

Polymethylmethacrylate (POL-ee-meth-uhl-meth-AK-rill-ate) is a clear thermoplastic resin used to make windshields, visors, coatings for baths, advertising signs, and contact lenses. It is also widely used in dentistry and medicine. A thermoplastic resin is one that becomes soft when heated and hard when cooled. It can be converted back and forth any number of times between the solid and liquid states by further heating and cooling.

Polymethylmethacrylate (PMMA) is more transparent than glass and six to seventeen times more resistant to breakage than glass. Another advantage it has over glass is that, if it does break, it falls apart into dull-edged pieces. PMMA is resistant to water, inorganic acids and bases, but is vulnerable to many organic solvents.

PMMA was first synthesized in 1928 in the laboratories of the German chemical firm Röhm and Haas. After five years of research, one of the firm's founders, Otto Röhm, found a way of manufacturing sheets of polymethylmethacrylate. He patented his invention and the company was soon producing the



Poly(methyl methacrylate).
Red atoms are oxygen; white
atoms are hydrogen; black
atoms are carbon. Gray stick
shows a double bond.

PUBLISHERS RESOURCE GROUP

product under the trade name of Plexiglas®. At about the same time, the product was being developed independently in the United States by the Dow chemical company, who sold its product under the trade name of Lucite®. Over time, a number of other chemical companies produced their own version of PMMA under a variety of trade names, including Acrylite®, Acrypet®, Crinothene®, Degalan®, Diakon®, Elvacite®, Kallocryl®, Metaplex®, Osteobond®, Paraglas®, Perspex®, Pontalite®, Sumipex®, Superacryl®, and Vedril®.

HOW IT IS MADE

The monomer from which polymethylmethacrylate is made is methyl methacrylate, $\text{CH}_2=\text{C}(\text{CH}_3)\text{COOCH}_3$. Methyl

Interesting Facts

- Polymethylmethacrylate exhibits a phenomenon known as total internal reflection. That term means that a light beam transmitted through a solid tube made of PMMA reflects off the inner surface of the tube. This property allows a light beam to be transmitted around corners and bends and out the end of a tube made of PMMA.
- A black-light reactive tattoo ink made from PMMA and microspheres of fluorescent dye is used on wildlife to track activities such as migration and growth patterns.
- The spectator shield in hockey arenas is made of PMMA.
- Plexiglas® was exhibited by Rohm & Haas at the World's Trade Fair in Paris in 1937. One of the exhibits was a transparent violin made from the product.

methacrylate is made in the reaction between acetone cyanohydrin ($(\text{CH}_3)_2\text{COHCN}$) and methanol (methyl alcohol; CH_3OH) in the presence of a sulfuric acid catalyst. As with all polymers, methyl methacrylate can be polymerized by a number of agents, including heat, radiation, and certain chemicals known as free-radical initiators. Once polymerization of methyl methacrylate begins, it can be processed in a number of ways. One of the most common processing system used involves passing the liquid material between two polished stainless steel belts. The distance between the belts is set to the desired thickness for the acrylic sheet. The acrylic is cured by a process of cooling and heating after leaving the steel belt. The final product is then cut to desired lengths at the end of the production line.

COMMON USES AND POTENTIAL HAZARDS

The most common use for PMMA is as a glass substitute. PMMA offers many benefits over glass because it is more transparent, less dense, stronger, and shatterproof. In ophthalmology (the branch of medicine that deals with diseases of the eye and their treatment), PMMA is used to make replacement lenses for the eye when the original lens has been removed for

Words to Know

CATALYST a material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

MONOMER one of the small, relatively simple molecules from which polymers are made.

some reason, such as the growth of a cataract. It is also used to make hard contact lenses. Other medical and dental applications of the material include its use as a dental cement, for the manufacture of bases and linings for dentures (false teeth), and to make a bone cement used in the reconstruction of broken or damaged bones and to fix implants in place.

Powdered polymethylmethacrylate presents health hazards because it may cause irritation of the skin, eyes, and respiratory tract. This hazard is of concern to people who work with the material in its raw form.

The glass-like material with which most people come into contact poses no health hazard for humans.

FOR FURTHER INFORMATION

"Material Information Polymethylmethacrylate, PMMA, Acrylic." Goodfellow.

<http://www.goodfellow.com/csp/active/static/E/ME30.HTML>
(accessed on October 24, 2005).

Meikle, J. L. *American Plastic: A Cultural History*. New Brunswick, N.J.: Rutgers University Press, 1997.

"More on the Manufacturing of Acrylic Sheet." *Plastics Distributor & Fabricator Magazine* (November/December 2000).

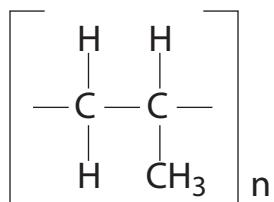
Available online at
<http://www.plasticsmag.com/features.asp?fIssue=Nov/Dec 00&aid=3053> (accessed on October 24, 2005).

"Polymethylmethacrylate." Kids' Macrogalleria. University of Southern Mississippi, The Polymer Science Learning Center.

<http://www.psfc.ws/macrog/kidsmac/pmma.htm> (accessed on October 24, 2005).

"Plexiglas® Primer." Ridout Plastics.

<http://www.ridoutplastics.com/plexprim.html> (accessed on October 24, 2005).



OTHER NAMES:

Propylene polymer;
1-propene
homopolymer

FORMULA:

$-\text{CH}(\text{CH}_3)\text{CH}_2-$

ELEMENTS:

Carbon, hydrogen

COMPOUND TYPE:

Organic polymer

STATE:

Solid

MOLECULAR WEIGHT:

Very large; 40,000 g/mol and up

MELTING POINT:

Varies: 165°C-170°C
(330°F-340°F)

BOILING POINT:

Not applicable

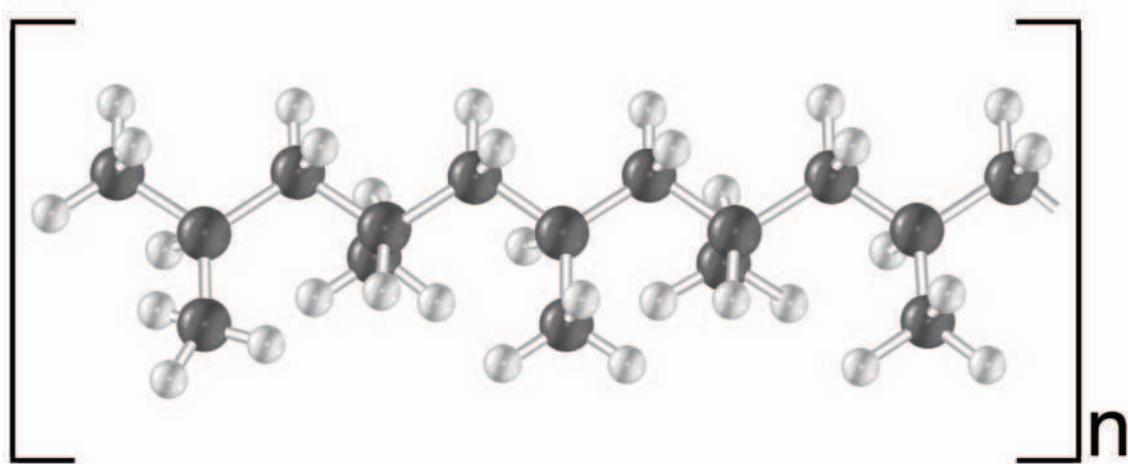
SOLUBILITY:

Insoluble in water and cold organic solvents;
softens in warm organic solvents, but does not dissolve;
soluble in hydrocarbons and halogenated hydrocarbons

KEY FACTS

OVERVIEW

Polypropylene (pol-ee-PRO-pih-leen) is a translucent white solid that is resistant to attack by heat, abrasion, inorganic acids and bases, most organic materials, and bacteria and fungi. It has high electrical resistance and tensile strength (its resistance to being pulled apart) and is very flexible. It takes on color well and can be coated with chrome. It can be prepared in a number of shapes and forms by extrusion and molding. Polypropylene is a thermoplastic polymer, meaning that it can be heated and cooled repeatedly, changing from a solid to a liquid and back again. Polypropylene is currently available under a number of trade names, including Amco®, Amerfil®, Azdel®, Beamette®, Clysar®, Daplen®, Dexon®, Epolene®, Gerfil®, Herculon®, Lambeth®, Lupareen®, Meraklon®, Mitsui Polypro®, Noblen®, Novolen®, Pellon®, Polypro®, Profax®, Propathene®, Propolin®, Propophane®, Shoallomer®, and Tuff-Lite®. Considerable dispute exists as to who should receive credit for inventing propylene. According to one history of the compound, it



Polypropylene. White atoms are hydrogen and black atoms are carbon. PUBLISHERS RESOURCE GROUP

was discovered independently about nine times. Patent disputes over the discovery lasted from the 1950s to 1989, when official credit was finally given to two researchers at the Phillips Petroleum Company, J. Paul Hogan (1919-) and Robert Banks (1921-1989). Phillips began selling polypropylene in 1951 under the trade name of Marlex®.

Much of the early research on polypropylene was conducted by the Italian chemist Giulio Natta (1903-1979), then an employee at the Italian chemical firm of Montecatini. While working earlier on the development of a related compound, polyethylene, Natta discovered some of the fundamental principles that governed the successful commercial production of polymers. In 1957, Montecatini began producing its own version of polypropylene. Six years later, Natta and his colleague, German chemist Karl Ziegler (1898-1973) shared the Nobel Prize in chemistry for their research on polymers.

HOW IT IS MADE

Polypropylene is made by the polymerization of propylene (propene; $\text{CH}_3\text{CH}=\text{CH}_2$). Polymerization is the process by which a single monomer unit (propylene in this case) is added to a second monomer of the same kind. The procedure is then repeated over and over again. Each time another monomer is added to the growing chain, the molecule gets larger and larger. Normally, polymerization is initiated by any of a number of agents, including radiation, light, or heat.

Interesting Facts

- Natta applied for a patent for polypropylene before telling Ziegler of his success in making the compound. Ziegler was so angry that the two men did not speak for many years. They reconciled only when they were both awarded the Nobel Prize at the same time.
- Australian bank notes (paper money) are made from polypropylene, which makes them more durable than paper bills.

Polymerization of propylene presents a somewhat different problem, however, because of the presence of methyl (-CH_3) groups extending off the main chain of the molecule ($[\text{-CH}(\text{CH}_3)\text{CH}_2]_n$). If polymerization is allowed to proceed on its own, some methyl groups will extend in one direction from the main chain, and others in a different direction. The product of this reaction is an amorphous product, one without crystalline shape, that has only a few very limited uses. To produce crystalline polypropylene, with all the desirable properties noted above, polymerization must be controlled to make sure that all methyl groups are on the same side of the main chain. One of Natta and Ziegler's great contributions was the discovery of catalysts capable of achieving the correct orientation of methyl groups. They found that metal halides, such as titanium chloride, could produce this effect. More recently, the German chemist Walter Kaminsky (1941-) and his colleagues found another group of catalysts that could polymerize crystalline polypropylene even more efficiently, a group of compounds called metallocenes. The manufacture of polypropylene today depends heavily on the use of such catalysts.

COMMON USES AND POTENTIAL HAZARDS

About one-third of all the polypropylene consumed in the United States is used to make fibers, for the manufacture of products such as blankets, fabrics, carpets, yarns, fish nets, protective clothing, laundry bags, and ropes. The next largest

Words to Know

CATALYST a material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

POLYMER a compound consisting of very large molecules made of one or two small repeated units called monomers.

uses are in the production of rigid packaging materials, such as crates, food containers, and bottles; in household products, such as dishes, bowls, outdoor carpeting, and outdoor furniture; and in packaging film. Hospitals use many surgical objects made out of polypropylene, taking advantage of its low cost and ability to be sterilized. Automobile manufacturers use the compound almost everywhere on the body of their cars. About 20 percent of all the polypropylene produced is used to make a large variety of products, including wire and cable insulation, medical tubing, pipe fittings, battery cases, drinking straws, and packaging foam.

No human health hazards have been identified for polypropylene in the form in which most people come into contact with the compound.

FOR FURTHER INFORMATION

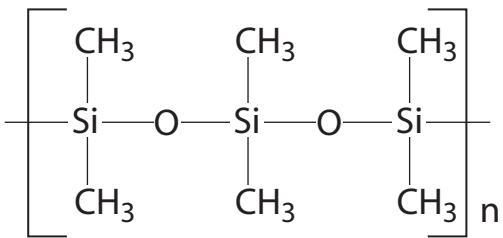
Karger Kocsis, J. *Polypropylene An A Z Reference*. New York: Springer Verlag, 1998.

Meikle, Jeffrey L. *American Plastic: A Cultural History*. Piscataway, N.J.: Rutgers University Press, 1997.

“Polypropylene.” The Macrogallery. Polymer Learning Center, University of Southern Mississippi.
<http://www.pscl.ws/macrogcss/pp.html> (accessed on October 24, 2005).

“Polypropylene.” *World of Chemistry*. Edited by Robyn V. Young. Detroit, Mich.: Gale, 1999.

See Also Polyethylene



OTHER NAMES:

Siloxane;
organosiloxane;
silicone

FORMULA:

Varies

ELEMENTS:

Varies: Silicon
(always), oxygen
(always), carbon
(almost always),
hydrogen (almost
always)

COMPOUND TYPE:
Inorganic polymer

STATE:
Solid or liquid

MOLECULAR WEIGHT:
Varies

MELTING POINT:
Varies

BOILING POINT:
Not applicable

SOLUBILITY:
Varies

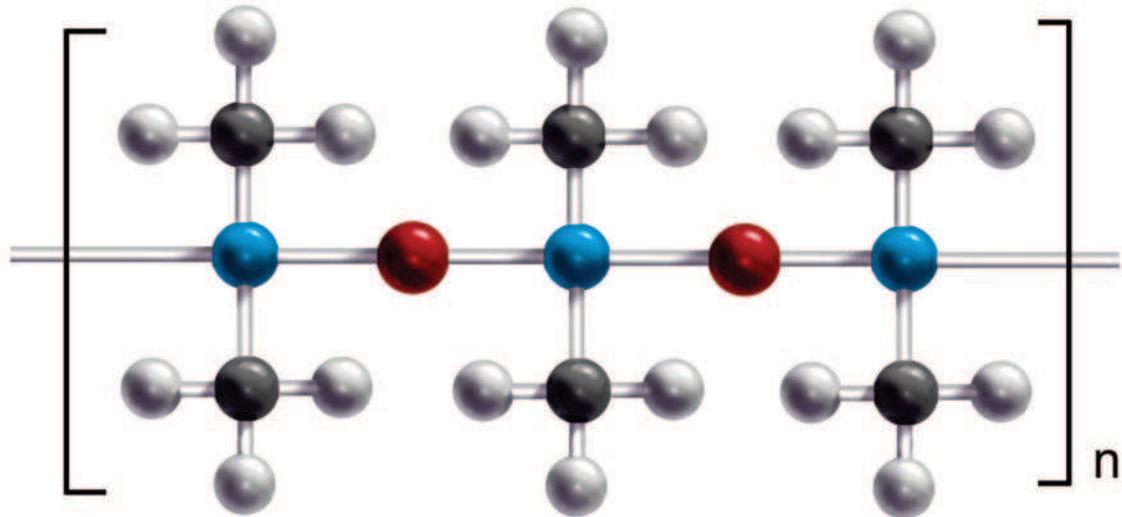
KEY FACTS

Polysiloxane

OVERVIEW

The term polysiloxane (pol-ee-sill-OK-sane) refers to a class of compounds whose molecules consist of a silicon-oxygen backbone ($-\text{Si-O-Si-O-Si-O-Si}-$)_n arranged either in a linear or cyclic (ring) pattern. Each silicon in the chain has two additional oxygen atoms attached to it. In many cases, the polysiloxanes also have one or more alkyl groups attached to the main chain replacing one or more of the oxygens. An alkyl group is an alkane, a saturated hydrocarbon, lacking one hydrogen atom. Examples of alkyl groups are the methyl ($-\text{CH}_3$) and ethyl ($-\text{CH}_2\text{CH}_3$) groups. In one common type of siloxane, all of the oxygens that are not a part of the backbone of the chain are replaced by methyl groups. Polysiloxanes that contain alkyl groups are known as organosiloxanes or, more commonly, silicones.

Researchers have now learned how to modify organosiloxane polymers by using various alkyl groups in chains of various lengths and conformations to produce a very wide array of products. All are organosiloxanes, but with very



Polysiloxane. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and turquoise atoms are silicon. PUBLISHERS RESOURCE GROUP

different physical and chemical properties and, hence, with very different applications. They range from liquid to gel to a semi-solid, rubber-like material. In general, the organosiloxanes tend to be chemically inert and resistant to heat. Although not identical, the terms polysiloxane, siloxane, and organosiloxane, and silicone are often used interchangeably.

The chemistry of silicon has long fascinated chemists. The element lies just below carbon in the periodic table, meaning that its physical and chemical properties are similar to carbon from which millions of organic compounds are made. The hope has long been to discover if silicon also can make a diverse number of compounds, as carbon can. One researcher who pursued this question in the early twentieth century was the English chemist Frederic Stanley Kipping (1863-1949). Kipping made use of a new kind of chemical reaction, called the Grignard reaction, after the French chemist Victor Grignard (1871-1935) who invented it. With the Grignard reaction, Kipping was able to make a number of silicon compounds to which were bonded alkyl groups in a variety of conformations. These compounds were the first organosiloxanes produced and studied in any detail, earning Kipping the title of "Father of Silicone Chemistry." During his lifetime, Kipping wrote more than 50 scholarly papers on his work. He did not, however, see the possibility of commercial applications for his discoveries.

The commercialization of silicones did not become possible until the early 1940s when a research chemist at the General Electric Company, E. G. Rochow (1909–2002) found an efficient way of making organosiloxanes in large quantities. Rochow's discovery came at an opportune time, the beginning of World War II. A number of military applications were found for the new product almost immediately. Within five years, a number of major chemical companies, including General Electric, Dow-Corning, Union Carbide, Stauffer Chemical, Wacker-Chemie, and Farbenfabriken Bayer A. G., had begun making organosiloxanes in large quantities.

HOW IT IS MADE

A variety of production methods is available for making different types of organosiloxanes. Probably the most popular approach begins with the reduction of silicon dioxide (sand; SiO_2) in an electric furnace. In that reaction, the silicon dioxide breaks down into silicon and oxygen: $\text{SiO}_2 \rightarrow \text{Si} + \text{O}_2$. The silicon is then treated with methyl chloride (CH_3Cl), which produces a mixture of methylchlorosilanes. Methylchlorosilanes are compounds that contain a single silicon atom to which are attached one or more methyl groups and one or more chlorine atoms. The methylchlorosilane produced in largest amount is dimethyldichlorosilane, $(\text{CH}_3)_2\text{SiCl}_2$. When this compound is treated with water, all of the chlorines are replaced by oxygens and the resulting compound polymerizes spontaneously. That is, individual molecules of the newly-formed silicon-carbon-oxygen compound begin to react with each other to form long organosiloxane chains. The size and character of the chain can be controlled by adding compounds that stop the polymerization reaction at some given point, resulting in the formation of a liquid, a gel, a semi-solid, or some other form of the product.

COMMON USES AND POTENTIAL HAZARDS

Organosiloxanes are very heat resistant, so they do not easily melt, like most other organic compounds. They are also water-repellant and can withstand extremes of sunlight, moisture, cold, and attack by most chemicals. These properties make them useful for protective coatings, electrical insulation, adhesives, lubricants, paints, and rubber-like materials. Some silicones are also used to make nonstick

Interesting Facts

- One of the most famous footprints in the world—that made by Neil Armstrong during his landing on the Moon in 1969—was made with a boot with a silicone rubber sole.
- Silicone was first used for breast implants in the 1960s for women who had undergone mastectomies, surgical removal of their breasts. Silicone implants later became popular with women who had no medical problems, but wanted larger breasts. By the 1980s, many women with breast implants

began complaining of pain, chronic fatigue, inflammation of breast tissue, and other medical problems thought to be associated with their silicone breast implants. Some of these women sued Bristol-Meyers Squibb, Dow Chemical, 3M, and other companies who made silicone for breast implants. By 1995, more than 20,000 legal claims had been brought against Dow alone. Faced with this staggering number of claims, the company went bankrupt.

surfaces, such as pans and spatulas for the kitchen. They are also used in making other kitchen items, such as oven mitts, because of their heat resistance.

Methyl silicones are also a major ingredient in personal care products. They are added to shaving lotions to provide lubrication and to give these products a non-greasy, yet silky, feeling. They help hair-styling products to spread more easily, and they increase the skin protection factor (SPF) in sunscreens. Silicones are also used in deodorants, perfumes, and nail polishes.

Some industrial and commercial applications of silicones include:

- As adhesives and sealants in many aircraft parts, including doors, windows, wings, fuel tanks, hydraulic switches, overhead bins, wing edges, leading gear electrical devices, vent ducts, engine gaskets, electrical wires and black boxes;

Words to Know

ALKANE a hydrocarbon where all the bonds between atoms are single bonds, the carbons are linked in a chain (except for methane, CH₄), and every carbon atom is saturated, or bonded to the maximum number of hydrogen atoms.

ALKYL GROUP an alkane that is lacking one hydrogen atom.

HYDROCARBON a chemical that consists of carbon and hydrogen.

- In the construction business as sealants for all kinds of building materials, including concrete, glass, granite, steel and plastic;
- In the manufacture of many automobile parts, including air bags, gaskets, headlamps, hydraulic bearings, ignition cables, radiator seals and hoses, shock absorbers, spark plug boots and ventilation flaps;
- As heat transfer agents;
- For the weatherproofing of concrete and other surfaces;
- In the production of surgical membranes and implants;
- In a great variety of electrical and electronic appliances, devices, and parts, including computer cases, keyboards, copy-machine components, and telephones; and
- For providing attractive and sturdy finishes for fabric and clothing.

FOR FURTHER INFORMATION

"The Basics of Silicon Chemistry." Dow Chemical Company.

<http://www.dowcorning.com/content/sitech/sitechbasics/silicones.asp> (accessed on October 26, 2005).

"Heat and Chemical Resistant Silicone Rubber. Silicones 2. Organic Silicon Chemistry." ChemCases.com Kennesaw State University.
<http://www.chemcases.com/silicon/silicone.htm> (accessed on October 26, 2005).

"Silicones." Macrogalleria, University of Southern Mississippi.
<http://www.psfc.ws/macrogcss/silicone.html> (accessed on October 26, 2005).

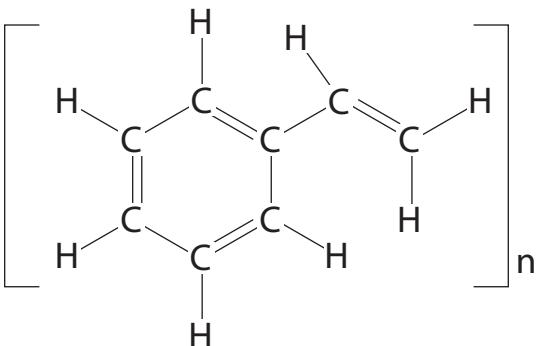
POLYSILOXANE

“Silicones Science On Line.” Centre Européen des Silicones.

<http://www.silicones science.com/> (accessed on October 26, 2005).

“What Are Silicones?” Silicones Environmental, Health and Safety Council of North America.

<http://www.sehsc.com/index.asp> (accessed on October 26, 2005).



OTHER NAMES:

Styrofoam

FORMULA:

-[-CH₂C₆H₅]_n

ELEMENTS:

Carbon, hydrogen

COMPOUND TYPE:

Organic polymer

STATE:

Solid

MOLECULAR WEIGHT:

Varies g/mol

MELTING POINT:

Varies widely; ranges from 190°C-260°C (370°F-500°F)

BOILING POINT:

Not applicable

SOLUBILITY:

Insoluble in water and inorganic acids and bases; soluble in many organic solvents, including ethylbenzene, chloroform, carbon tetrachloride, and tetrahydrofuran

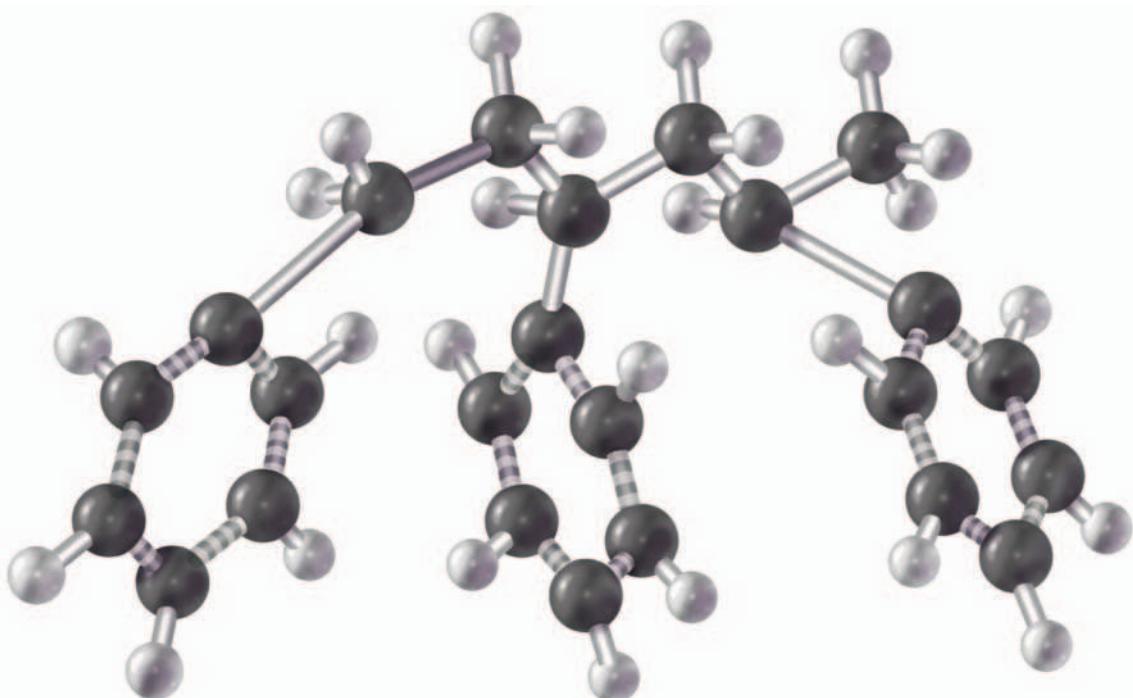
K
E
Y
F
A
C
T
S

OVERVIEW

Polystyrene (pol-ee-STYE-reen) is a thermoplastic polymer made from styrene. A thermoplastic polymer is a material that can be repeatedly softened and hardened by alternately heating and cooling. Styrene is a hydrocarbon derived from petroleum with the formula C₆H₅CH=CH₂. The presence of the double bond in the styrene molecule makes it possible for styrene molecules to react with each other in long chains that constitute the polymer polystyrene.

Polystyrene is a hard, strong, transparent solid highly resistant to mechanical impact. It is an excellent thermal (heat) and electrical insulator, is easily shaped and molded in the liquid state, and takes dyes readily. It can be produced in a wide variety of shapes and forms, including sheets, plates, rods, beads, and foams.

The history of polystyrene dates to 1839 when a German apothecary (druggist) named Eduard Simon discovered styrene in petroleum. Later scientists attempted to incorporate



Polystyrene. White atoms are hydrogen and black atoms are carbon. Gray sticks indicate double bonds. Striped sticks indicate benzene rings.

PUBLISHERS RESOURCE GROUP

styrene into some of the commercial products they made, such as rubber tires. But a polymer based on styrene was not actually produced until 1930 when researchers at the German chemical firm of I. G. Farben discovered how to make the material. Farber's parent company, BASF, shortly made the product commercially available and in 1937, Dow Chemical first made the compound available in the United States. During World War II, polystyrene was used for the manufacture of synthetic rubber products. After the war, the number of commercial and industrial uses expanded rapidly. Today, it is virtually impossible to avoid polystyrene products in one's daily life.

HOW IT IS MADE

Compounds like styrene with double bonds often polymerize easily. The double bond on a styrene molecule breaks open and a hydrogen atom from a second styrene molecule adds to one side of the double bond, while the rest of the second styrene molecule adds to the second side of the double bond. The product of this reaction still has a double bond.

Interesting Facts

- Polystyrene is sold commercially under more than a hundred trade names, the most famous of which is probably Styrofoam®.
- Only about 5 percent of a styrofoam cup is polystyrene. The rest is air.
- One of the innovative uses for polystyrene is as a building material for the construction of new houses. Scientists suggest

that it is perfect for the purpose: lightweight, inexpensive, strong, a good insulator, and available all over the world. One of the first applications suggested for polystyrene as a building material is in the construction of houses in Afghanistan, where many families have lost their homes after two decades of wars and earthquakes.

So the reaction can be repeated a second time; and a third time; and a fourth time; and so on. One goal of research on polystyrene has been to determine how the size of the polystyrene affects its properties (and, therefore, its uses) and how to stop the polymerization reaction at some desired point.

All that is needed to start the polymerization of styrene is a material that will cause the first double bond to break. Such materials are known as polymerization initiators. One of the most common initiators used in the polymerization of styrene is benzoyl peroxide ($C_6H_5COOOCOC_6H_5$). Once the polymerization reaction begins, it tends to release enough energy for the reaction to continue on its own.

An especially popular form of polystyrene is known as expanded polystyrene. It is made by blending air with molten polystyrene to make a lightweight foam sold under the trade name of Styrofoam®.

COMMON USES AND POTENTIAL HAZARDS

Polystyrene is the fourth largest thermoplastic polymer made in the United States by production volume. It

is used in the manufacture of hundreds of commercial, industrial, household, and personal articles. Some examples include:

- Plastic model kits and toys;
- Containers with lids; disposable cups, plates, knives, forks, and spoons;
- "Jewel" cases for compact discs and cases for audiocassettes;
- Plastic coat hangers and plastic trays;
- Refrigerator doors and air conditioner cases;
- Housing for machines; and
- Cabinets for clocks, radios, and television sets.

Some uses of expanded polystyrene include:

- In all kinds of containers to keep foods either hot or cold (such as ice chests);
- Egg cartons;
- Fillers in shipping containers;
- Packages for carry-out foods;
- Insulation for buildings;
- In the construction of boats; and
- For the construction of some types of furniture.

Polystyrene dust and powder formed during production can be a mild irritant to the eyes, skin, and respiratory system. But even for workers in the field, the risk is regarded as being very low. A more serious problem posed by the compound is the risk it poses for the environment. About half of all the polystyrene produced in the United States is used for packaging and "one-time use" purposes. That is, someone uses the product and then throws it away. Since polystyrene does not readily decompose, it tends to accumulate in landfills and dumps. Some environmentalists point out that large volumes of discarded polystyrene contribute significantly to the nation's solid waste disposal problems. Industry spokespersons, however, point out that polystyrene accounts for less than one percent of all solid wastes. In any case, a number of industries and companies have attempted to reduce the amount of polystyrene used in their products in order to

Words to Know

HYDROCARBON a compound consisting of carbon and hydrogen.

POLYMER a compound consisting of very large molecules made of one or two small repeated units called monomers.

THERMOPLASTIC can be repeatedly softened and hardened by alternately heating and cooling.

cut back on their contribution to the solid waste disposal problem.

FOR FURTHER INFORMATION

Boyd, Clark. "Polystyrene homes planned for Afghans." BBC News.
<http://news.bbc.co.uk/2/hi/technology/3528716.stm> (accessed on October 26, 2005).

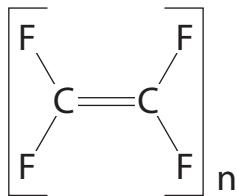
"Energy & Waste Landfilling." Energy Information Administration, Department of Energy.
<http://www.eia.doe.gov/kids/energyfacts/saving/recycling/solidwaste/landfiller.html> (accessed on October 26, 2005).

"Polystyrene." U.S. Environmental Protection Agency, Emissions Factors and Policy Applications Center.
http://www.epa.gov/ttn/chief/ap42/cho6/final/co6so6_3.pdf (accessed on October 26, 2005).

"Polystyrene Packaging Delivers!" Polystyrene Packaging Council.
<http://www.polystyrene.org/> (accessed on October 26, 2005).

Sims, Judith. "Polystyrene." In *Environmental Encyclopedia*. 3rd ed. Edited by Marci Bortman and Peter Brimblecombe. Detroit, Mich.: Gale, 2003.

See Also Styrene



OTHER NAMES:
See Overview.

FORMULA:
 $-\text{CF}_2\text{-}\text{I}_n$

ELEMENTS:
Carbon, fluorine

COMPOUND TYPE:
Organic polymer

STATE:
Solid

MOLECULAR WEIGHT:
Varies

MELTING POINT:
Varies; most common range: 302°C - 310°C (575°F - 590°F)

BOILING POINT:
Not applicable

SOLUBILITY:
Insoluble in water, concentrated acids, other inorganic solvents and almost all organic solvents

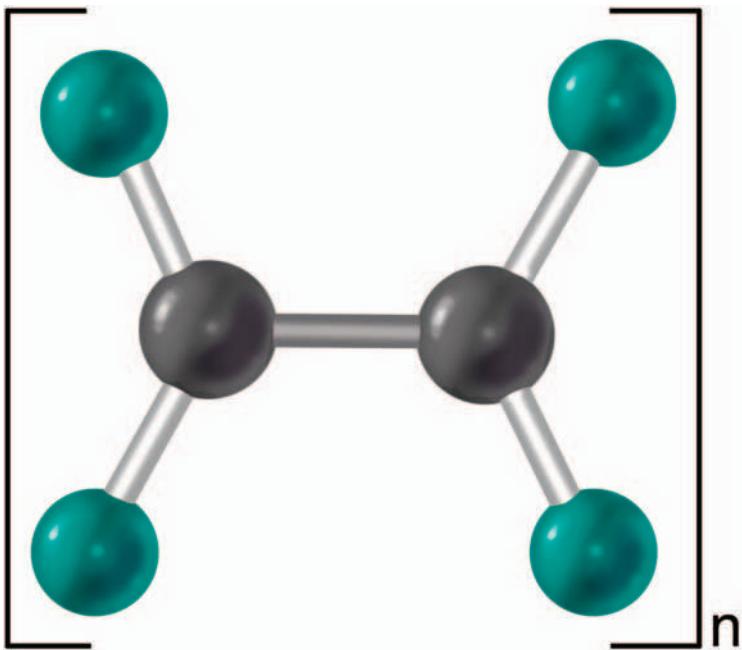
KEY FACTS

OVERVIEW

Polytetrafluoroethylene (POL-ee-tet-ruh-FLUR-oh-ETH-eh-leen) is also known as polytetrafluoroethene, tetrafluoroethylene polymer, PTFE, and Teflon®. Polytetrafluoroethylenes are thermoplastic polymers made from the monomer tetrafluoroethylene ($\text{CF}_2=\text{CF}_2$). A thermoplastic polymer is a material that can be repeatedly softened and hardened by alternately heating and cooling. The polytetrafluoroethylenes are well known by the trade name of Teflon® although they are also available under more than a hundred other trade names, including Aflon®, Algloflon®, Ethicon®, Fluon®, Ftorlon®, Halon®, Molykote®, Polyflon®, Polytet®, and PTFE®. DuPont Chemical, one of the two largest manufacturers of polytetrafluoroethylene, makes at least ten grades of Teflon®. These products differ from each other in the physical form in which they are provided (powders, aqueous dispersions, yarn, or film, for example) and size (molecular weight) of the product (ranging from low-molecular weight to high-molecular weight).

POLYTETRAFLUOROETHYLENE

Polytetrafluoroethylene.
Black atoms are carbon;
turquoise atoms are fluorine.
Gray stick indicates double
bond. PUBLISHERS RESOURCE
GROUP



Polytetrafluoroethylene has the lowest coefficient of friction of any known substance. The coefficient of friction is a measure of how easily one substance slides over the surface of a second substance. Polytetrafluoroethylene's low coefficient of friction means that nothing will stick very well to its surface, accounting for Teflon®'s primary use in the manufacture of non-stick products.

Polytetrafluoroethylene was invented in 1938 by Roy J. Plunkett (1910-1994) quite by accident. As a research chemist at DuPont's Jackson Laboratory, Plunkett was studying compounds that might be used for refrigerants. He kept the compounds in steel tanks and was surprised on one occasion to find that the gas he wanted did not leave the storage tank when the valve was opened. He cut the tank open to see what had happened to the gas and found a waxy white material. Upon analysis, the material turned out to be polytetrafluoroethylene. The gas stored in the tank, the potential refrigerant, was tetrafluoroethylene. It had undergone polymerization spontaneously within the tank, making it possible for Plunkett to discover one of the most remarkable synthetic products in the world.

Interesting Facts

- Because Teflon® does not stick to anything else, cookware manufacturers must use a special process to get it to stay on pots and pans. They sometimes begin by blasting the pan with sand or grit to roughen the surface. Then they apply a special primer that makes the Teflon® adhere to the pan's surface.
- Teflon®'s non-stick quality has been compared to trying to make one piece of ice stick to another piece of ice.

HOW IT IS MADE

Molecules of tetrafluoroethylene contain double bonds. Any compound with double bonds has the ability to form polymers. Polymerization of tetrafluoroethylene occurs when the double bond in one molecule breaks apart. A fluorine atom from a second molecule of the monomer then adds on to one end of the broken double bond. The rest of the second molecule adds to the other end of the broken double bond. A "double-molecule," consisting of two monomers joined to each other forms: $\text{CF}_2=\text{CF}_2 + \text{CF}_2=\text{CF}_2 \rightarrow \text{CF}_3\text{CF}_2\text{CF}_2=\text{CF}_2$. The product of this reaction also contains a double bond. So the process can be repeated to form another product consisting of three monomer molecules: $\text{CF}_3\text{CF}_2\text{CF}_2=\text{CF}_2 + \text{CF}_2=\text{CF}_2 \rightarrow \text{CF}_3\text{CF}_2\text{CF}_2\text{CF}_2\text{CF}_2=\text{CF}_2$. The process is repeated hundreds or thousands of times producing a long chain of monomers with the general formula $[-\text{CF}_2-]_n$.

COMMON USES AND POTENTIAL HAZARDS

Perhaps the best known application of polytetrafluoroethylene is in kitchen utensils with non-stick coatings, such as pots, pans, and spatulas. Polytetrafluoroethylene is also used to coat fibers to make them water-repellant and stain-resistant. Water will bead up and roll off the surface of clothing and other materials coated with polytetrafluoroethylene instead of penetrating the fabric and possibly

leaving a stain. Polytetrafluoroethylene is available as a spray treatment for carpets and furniture, forming a molecular shield to prevent water or oil-based stains from penetrating the material. Some carpets come pre-treated with a polytetrafluoroethylene product to keep them clean and fresh. The compound can also be used on wood and plastic flooring to protect it from dirt, stains, and moisture.

Automobile manufacturers used polytetrafluoroethylene in a variety of ways. Windshield wiper blades coated with polytetrafluoroethylene are smoother and stronger than uncoated blades. Automobile paint can be coated with polytetrafluoroethylene to protect a car's finish from tree sap, insects, and other residues. Automobile upholstery is often treated with polytetrafluoroethylene to protect against stains caused by spilled drinks and dirty shoes. Polytetrafluoroethylene added to oil makes it flow through an engine more smoothly, reducing wear and tear on the engine.

Polytetrafluoroethylene is used widely for a number of industrial applications. Some armor-piercing bullets are coated with the compound to reduce friction when the bullet leaves the gun barrel and increases the ease with which it breaks through armor. Many electrical cables are insulated with polytetrafluoroethylene, which is not combustible or conductive. Food processing equipment made with polytetrafluoroethylene is easier to clean and more efficient for cooking and baking than non-polytetrafluoroethylene equipment. Industrial bakers often use equipment coated with some type of polytetrafluoroethylene. The product can also be used to coat stainless steel, carbon steel, aluminum, steel alloys, brass, magnesium, glass, fiberglass, plastics, and rubber. Outdoor signs are sometimes coated with polytetrafluoroethylene to make them last longer and resist stains.

Polytetrafluoroethylene has long been regarded as an essentially safe compound with no known health effects on humans or experimental animals. Recently, questions have been raised about possible health hazards of one of the compounds used in the manufacture of polytetrafluoroethylene, perfluorooctanoic acid (PFOA). Some studies suggest that PFOA may be responsible for birth defects and the development of cancer in people who have been exposed to the chemical. Other studies show that 96 percent of the children tested in 23 states and the District of Columbia in 2001 had detectable levels of PFOA in their blood. Federal

Words to Know

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

THERMOPLASTIC Able to be repeatedly softened and hardened by alternately heating and cooling.

agencies have not yet confirmed the level of risk that PFOA poses, if any, and have not listed any other chemicals used in the manufacture of polytetrafluoroethylene as hazardous to human health.

FOR FURTHER INFORMATION

“PTFE Specifications.” Boedeker Plastics.

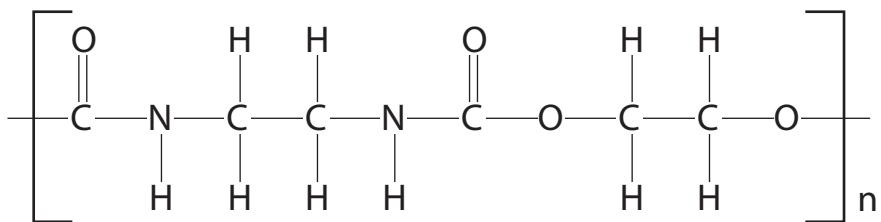
http://www.boedeker.com/ptfe_p.htm (accessed on October 26, 2005).

Summer, Chris. “Teflon’s Sticky Situation.” BBC News Online (October 7, 2004). Available online at <http://news.bbc.co.uk/1/hi/magazine/3697324.stm> (accessed on October 26, 2005).

“Technical Information: Teflon® Fluorocarbon Resin.” Omega. Stamford, Conn.: Omega Engineering, Inc. 2000. Available online at http://www.omega.com/pdf/tubing/technical_section/teflon_flourocarbon.asp (accessed on October 26, 2005).

“Teflon (PTFE: polytetrafluoroethylene).” Fluoride Action Network. http://www.fluoridealert.org/pesticides/polytetrafluoroethylene_page.htm (accessed on October 26, 2005).

See Also Polyethylene



OTHER NAMES:

None

FORMULA:

-[CONH-C₆H₄-NCOO-CH₂CH₂O-]_n; other structures are possible

ELEMENTS:

Carbon, hydrogen, oxygen, nitrogen

COMPOUND TYPE:

Organic polymer

STATE:

Solid

MOLECULAR WEIGHT:

Varies; very large

MELTING POINT:

Variable

BOILING POINT:

Not applicable

SOLUBILITY:

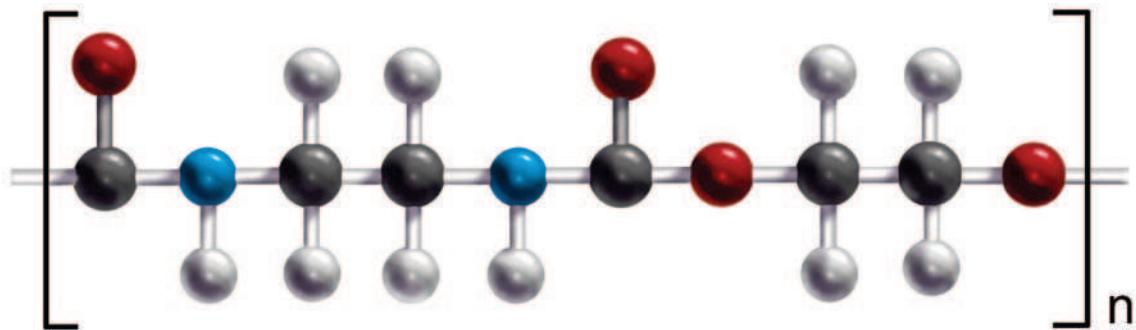
Insoluble in water; soluble in aromatic hydrocarbons, such as benzene and toluene

K E Y F A C T S

OVERVIEW

Polyurethanes (pol-ee-YUR-eth-anes) are a group of thermoplastic polymers formed in the reaction between a diisocyanate and a polyol, an alcohol with two or more hydroxyl (-OH) groups. Diisocyanates are compounds that contain two isocyanate (-N=C=O) groups.

Polyurethanes are available in a variety of forms, including fibers, foams, coatings, and elastomers, rubber-like materials. Each form of polyurethane has its own set of physical and chemical properties. For example, fibers are moisture proof, stretchable, and resistant to the flow of electric current. Foams can be either rigid or flexible, with densities as low as 32 kilograms per cubic meter (2 pounds per cubic foot) to as high as 800 kilograms per cubic meter (50 pounds per cubic foot). They are excellent thermal (heat) insulators. Polyurethane coatings are hard, glossy, flexible, readily adhesive to other surfaces, and resistant to abrasion, weathering, and most inorganic chemicals. Elastomeric polyurethanes are resistant to abrasion, weathering, and most organic solvents.



Polyurethane. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atoms are nitrogen. Gray sticks show double bonds. PUBLISHERS RESOURCE GROUP

With this variety of properties, polyurethanes have a very wide range of uses.

The basic process for making polyurethanes was first developed in 1937 by the German chemist Otto Bayer (1902–1982), who patented his discovery and founded a company for its commercial production. The polyurethanes were first put to widespread use during World War II as substitutes for natural rubber in the manufacture of tires for military uses. The first rigid polyurethane foam was made in 1940, again for military purposes; the first polyurethane adhesive, for the joining of rubber and glass, in 1941; and the first use of polyurethane as an insulator, in beer barrels, in 1948. Polyurethane fibers were also used during World War II for the manufacture of protective clothing to be worn by soldiers in case of attacks by poison gas. In the late 1950s, a stretchable material made of polyurethane called spandex was introduced.

HOW IT IS MADE

Polyurethanes are formed by a reaction known as rearrangement. In this reaction, a hydrogen atom from the polyol (alcohol with two or more hydroxyl, -OH, groups) leaves the polyol and attaches itself to one of the nitrogen atoms in the diisocyanate molecule. The remnant of the polyol left after the hydrogen atom leaves then attaches itself to the carbon atom next to the nitrogen that has just received the hydrogen. The end result of this reaction is that two molecules, a diisocyanate molecule and a polyol molecule, have joined to make a single molecule. But that molecule has the same general structure as the original

Interesting Facts

The U.S. polyurethane industry produced 2,915 million kilograms (6,393 million

pounds) of polyurethane in 2002.

reactants. So the reaction can be repeated with a second diisocyanate molecule and a second polyol molecule adding to the former product.

All that is needed to make this reaction happen is a suitable catalyst, a substance that will encourage the first hydrogen atom to leave the polyol and move to the diisocyanate. The catalyst most frequently used is diazobicyclo[2.2.2]octane, more commonly known as DABCO. DABCO provides an initial “tug” on the hydrogen atom that needs to be moved. Once it gets the reaction started, DABCO has no further role and can be recovered for re-use.

Various forms of polyurethane are made by adding additional steps to this fundamental process. For example, polyurethane foams are made by adding water and carbon dioxide to the molten polymer. The carbon dioxide creates bubbles, which makes the mixture rise like bread and then harden into a foam. Flexible polyurethane for fibers and foam rubber can be made by adding polyethylene glycol, a softening agent, to the mixture.

COMMON USES AND POTENTIAL HAZARDS

Flexible and rigid foams are the most popular types of polyurethane made in the United States. Flexible foams are used for cushioning, as in mattresses, upholstered furniture, and automobile seats. Semiflexible foams are component of carpet pads, sponges, packaging, and door panels for automobiles. Rigid polyurethane foams are used to produce insulation for roofs, refrigerators, and freezers. The construction industry accounted for more than half of all the rigid polyurethane foam made in 2002, primarily for

Words to Know

POLYMER a compound consisting of very large molecules made of one or two small repeated units called monomers.

THERMOPLASTIC capable of being repeatedly softened and hardened by alternately heating and cooling.

roofing and wall insulation. The automotive industry was the second largest user of rigid foams, where polyurethane was used for floor cushions, headliners, and heating and air-conditioning systems.

Polyurethane elastomers are flexible and strong. They can be stretched to a significant amount before returning to their original length. They are also shock absorbent. These properties account for their uses in tires, shoe soles, skateboards, and inline skates. Polyurethane elastomers can also be spun into fibers which are then used to make light, flexible clothing. Spandex, one of the most popular fabrics made of polyurethane fibers, is both sturdy and flexible. It is used to make bathing suits and exercise clothing because it can be distorted without losing its original shape or tearing. Polyurethane is also one of the main ingredients in the fabric known as pleather, or plastic leather. This synthetic form of leather is less expensive to produce and easier to dye than real leather.

Polyurethane coatings are strong and durable. They can be found on surfaces that need to be flexible, abrasion resistant, and chemical resistant, such as dance floors and bowling alleys. Water-based coatings are often used on airplanes and other transportation equipment. Polyurethane powder coatings are also used on fluorescent lights, refrigerators, car wheels and trim, lawnmowers, patio furniture, and ornamental iron.

Dust formed by the breakdown of polyurethane products may be an irritant to skin, eyes, and the respiratory system. No serious health problems have been associated with such materials, however. Some of the raw materials used in the manufacture of the polyurethanes, such as the isocyanates, do pose health risks. These risks are of concern primarily to workers who come into contact with those raw materials.

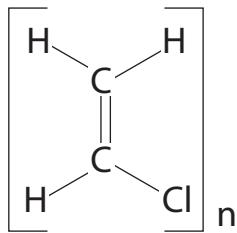
FOR FURTHER INFORMATION

Alliance for the Polyurethane Industry. "About Polyurethane"
<http://www.polyurethane.org/about/>(accessed on October 26, 2005).

"Making Polyurethanes." Polymer Science Learning Center, University of Southern Mississippi.
<http://www.psfc.ws/mactest/uresyn.htm> (accessed on October 26, 2005).

"Polyurethanes." Huntsman.
<http://www.huntsman.com/pu/> (accessed on October 26, 2005).

Vartan, Starre. "Pretty in Plastic: Pleather Is a Versatile, though Controversial, Alternative to Leather." *E* (September October 2002): 53 54.

**OTHER NAMES:**

See Overview.

FORMULA:

-[CH_2CHCl] $_n$

ELEMENTS:

Carbon, hydrogen, chlorine

COMPOUND TYPE:

Organic polymer

STATE:

Solid

MOLECULAR WEIGHT:

Varies

MELTING POINT:

Decomposes at 148°C
(298°F)

BOILING POINT:

Not applicable

SOLUBILITY:

Insoluble in water;
soluble in
tetrahydrofuran,
dimethylformamide,
dimethylsulfoxide

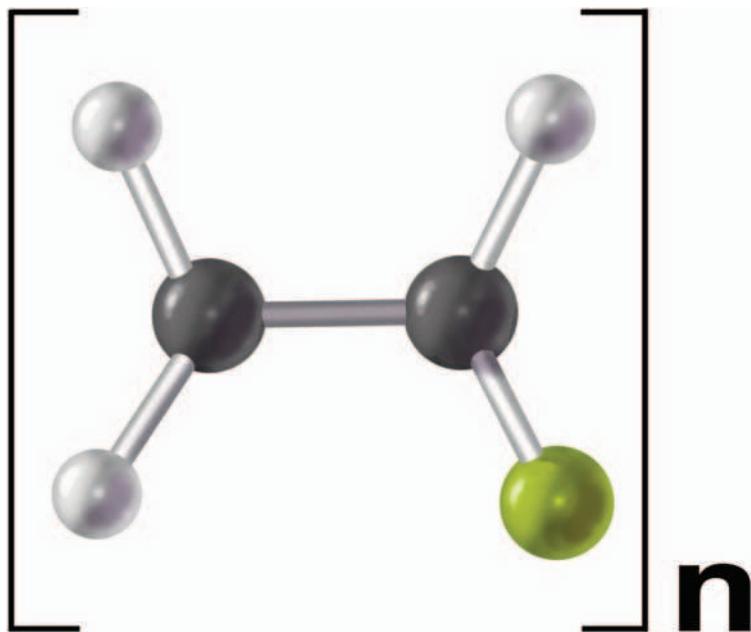
KEY FACTS**OVERVIEW**

Polyvinyl chloride (pol-ee-VYE-nul KLOR-ide) is also known as PVC, vinyl, chlorethylene homopolymer, and chloroethene homopolymer. It is the third most commonly produced plastic in the United States, exceeded only by polyethylene and polypropylene. It is offered commercially in a variety of formulations, usually as a white powder or colorless granules. The compound is resistant to moisture, weathering, most acids, fats and oils, many organic solvents, and attack by fungi. It is easily colored and manufactured in a variety of forms, including sheets, films, fibers, and foam.

Polyvinyl chloride was first discovered accidentally in 1835 by the French physicist and chemist Henry Victor Regnault (1810-1878). Regnault found that a container of gaseous vinyl chloride ($\text{CH}_2\text{CH=Cl}$) exposed to the sunlight gradually changed to a white powder. Regnault knew almost nothing about the composition of the powder or how it was formed. Polyvinyl chloride remained a subject of little or no interest to chemists for almost a century. German and

Poly(vinyl chloride). White atoms are hydrogen, black atoms are carbon and green atom is chlorine. Gray stick represents a double bond.

PUBLISHERS RESOURCE GROUP



Russian chemists made some efforts to find useful applications for the compound in the early twentieth century, without much success. The first patent for the production of the compound was awarded to the German chemist Friedrich Heinrich August Klatte (dates not available) in 1913, but Klatte never marketed the product for commercial use.

Then, in 1926, Waldo Lonsbury Semon (1898–1999), a chemist at B. F. Goodrich, rediscovered polyvinyl chloride, once again by accident. Semon had been assigned the task at Goodrich of finding a substitute for natural rubber as a lining for metal tanks and for finding a material that would bond the rubber substitute to the metal tank. One of the materials he studied was polyvinyl chloride. In his work, Semon did not so much rediscover polyvinyl chloride as to find new ways of working with the material. For example, he found that he could dissolve the compound in various organic solvents, converting them so that they could be molded, extruded, and formed. He also discovered that he could control the properties of the material by altering the amount of solvent used to dissolve powdered polyvinyl chloride.

Interesting Facts

- The recycling symbol for polyvinyl chloride is the number 3 inside a triangle made of three arrows.
- The environmental group Greenpeace has called for a global ban on the

production of polyvinyl chloride because the toxic substance dioxin is released (albeit, in very small amounts) as a byproduct of the compound's production.

Goodrich adapted Semon's discovery for two specific applications: shoe heels and the coating on chemical racks. Those applications were not profitable enough for Goodrich to continue making polyvinyl chloride. But Semon continued to look for new ways of adapting the compound for additional applications. He was eventually successful and by 1931 the company had begun to turn out a full line of polyvinyl chloride products in most of the forms currently available.

HOW IT IS MADE

Polyvinyl chloride is made by polymerizing vinyl chloride ($\text{CH}_2=\text{CHCl}$). Polymerization occurs when the double bond in vinyl chloride breaks, allowing one molecule of vinyl chloride to combine with a second molecule of vinyl chloride: $\text{CH}_2=\text{CHCl} + \text{CH}_2=\text{CHCl} \rightarrow \text{CH}_3\text{CHClCH}=\text{CHCl}$. The product of that reaction also contains a double bond, allowing the reaction to be repeated: $\text{CH}_3\text{CHClCH}=\text{CHCl} + \text{CH}_2=\text{CHCl} \rightarrow \text{CH}_3\text{CHClCH}_2\text{CHClCH}=\text{CHCl}$. Once again, the final product contains a double bond, and the reaction can be repeated again and again and again. The reaction is made possible by using some agent to cause double bonds to break. In the original experiments carried out by Regnault, Klatte, and others, that agent was sunlight. Chemists have long since learned, however, that a variety of chemicals known as peroxide initiators are more effective at breaking double bonds. Peroxide initiators are compounds with an oxygen-oxygen bond (-O-O-). One of the most widely used initiators is benzoyl peroxide ($\text{C}_6\text{H}_5\text{CO-O-O-C}_6\text{H}_5\text{CO}$).

COMMON USES AND POTENTIAL HAZARDS

An estimated 7 billion kilograms (15.7 billion pounds) of polyvinyl chloride were produced in the United States in 2006. It is available in more than 50 trade names, including Airex, Armodour, Astralon, Benvic, Bonloid, Chemosol, Chlorostop, Dacovin, Dorlyl, Flocor, Lucoflex, Norvinyl, Opanlon, Polivinit, Polytherm, Sicron, Takilon, Vinikulon, Viniplast, and Wilt Pruf. About three-quarters of that amount was made in a rigid format that is hard and inflexible. The remaining one-quarter was made in a flexible form, produced by adding materials to polyvinyl chloride that make it soft and pliable. About 75 percent of all rigid polyvinyl chloride (half of all the compound made in the United States) goes to the construction industry. It has replaced older building materials such as clay, concrete, and wood because it is inexpensive, lightweight, resistant to damage by the sun, and easy to assemble. The compound is used to make vinyl siding, windows, plumbing pipes, flooring, electric cables, roofing materials, and insulation for cables and wires.

Flexible polyvinyl chloride is used to make fibers and films for applications such as clothing, upholstery, plastic bottles, medical equipment, lightweight toys, shower curtains, and packaging films. Some of the medical equipment produced from polyvinyl chloride include bags to hold blood and other fluids, artificial heart valves, and tubes used in kidney dialysis. Medical products made from polyvinyl chloride are strong enough to be air-dropped to troops in combat zones. Automobile manufacturers use polyvinyl chloride in body side moldings, interior upholstery, engine wiring, floor mats, adhesives, dashboards, arm rests, and coatings under the vehicle.

The commercial and household products containing polyvinyl chloride are generally regarded as posing no threat to human health. However, a number of questions have been raised about possible health hazards and risks to the environment as a result of the process by which polyvinyl chloride is made, some of its applications, and its eventual disposal. For example, polyvinyl chloride is made from vinyl chloride, which itself is toxic and a carcinogen. People who work with vinyl chloride in production facilities are at risk for developing a form of liver cancer that may be related to exposure to vinyl chloride. Vinyl chloride, in turn, is made from chloride, a very toxic gas that poses health risks to people who work with it.

Some environmental health experts point out that products made of polyvinyl chloride may give off toxic or carcinogenic gases for short periods of time after they have been put into use. For example, shower curtains and automobile upholstery containing polyvinyl chlorides may release hazardous chemicals into the air for a few months after they are first installed.

Some concern has also been expressed about additives used with polyvinyl chloride. For example, some of the substances used to soften the compound belong to a family known as the phthalates, which are known carcinogens with toxic effects. Some health experts point out that infants may ingest small amounts of phthalates from PVC toys they chew on.

Finally, the disposal and destruction of compounds containing polyvinyl chloride may create environmental problems. Burning such products, for example, releases hydrogen chloride gas, a suffocating and toxic gas, into the atmosphere. Enough concern about the health and environmental hazards has arisen that some governmental bodies in Europe have placed limitations on the uses to which PVC products can be put.

FOR FURTHER INFORMATION

"Healthy Building Network."

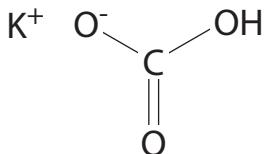
<http://www.healthybuilding.net/> (accessed on October 29, 2005).

Meikle, Jeffrey L. *American Plastic: A Cultural History*. Piscataway, N.J.: Rutgers University Press, 1997.

Thornton, Joe. "Environmental Impacts of Polyvinyl Chloride (PVC) Building Materials." A briefing paper for the U.S. Green Building Council, n.d. Available online at http://www.usgbc.org/Docs/LEED_tsac/PVC/CMPBS%20Original%20Submittal.pdf (accessed on October 29, 2005).

"Vinyl The Material." The Vinyl Institute.

<http://www.vinylinfo.org/materialvinyl/history.html> (accessed on October 29, 2005).



OTHER NAMES:	
Potassium acid carbonate; potassium hydrogen carbonate	
FORMULA:	
KHCO_3	
ELEMENTS:	
Potassium, hydrogen, carbon, oxygen	
COMPOUND TYPE:	
Acid salt (inorganic)	
STATE:	
Solid	
MOLECULAR WEIGHT:	
100.12 g/mol	
MELTING POINT:	
decomposes above 100°C (212°F)	
BOILING POINT:	
Not applicable	
SOLUBILITY:	
Soluble in water; insoluble in ethyl alcohol	

KEY FACTS

Potassium Bicarbonate

OVERVIEW

Potassium bicarbonate (poe-TAS-ee-yum buy-KAR-bo-nate) is a colorless crystalline solid or white powder with no odor and a salty taste. It occurs naturally in salt beds, sea water, silicate rocks, and a number of foods, primarily fruits and vegetables. Potassium bicarbonate is also present in the tissues of humans and other animals, where it is involved in a number of essential biological processes, including digestion, muscle contraction, and heartbeat. It is used primarily in cooking and baking, as a food additive, and in fire extinguishers.

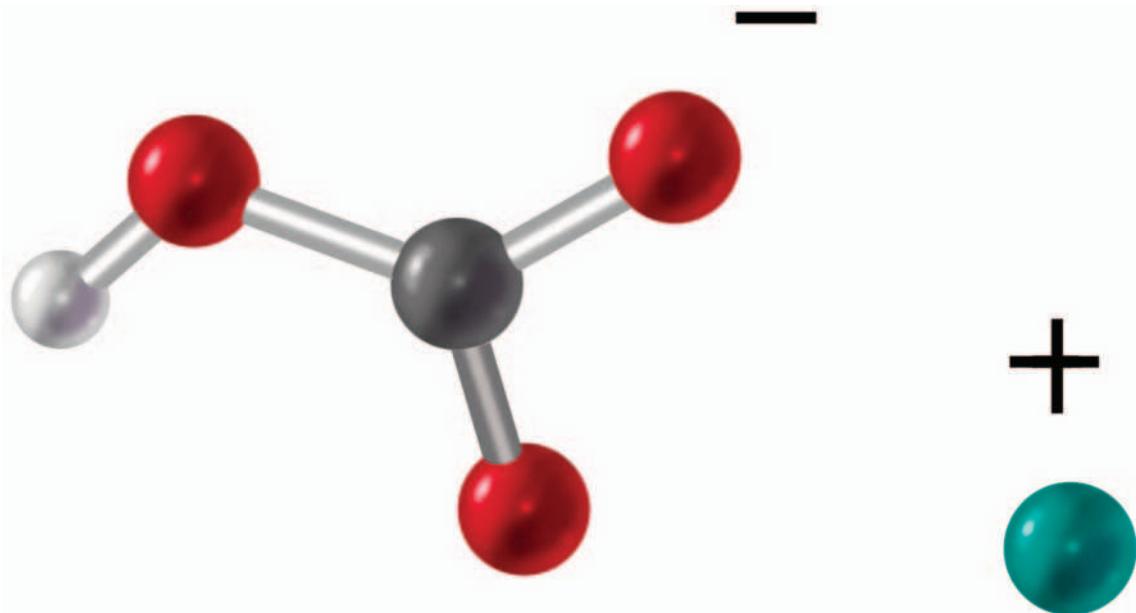
HOW IT IS MADE

Potassium bicarbonate is made by passing carbon dioxide gas through an aqueous solution of potassium carbonate: $\text{K}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O} \rightarrow 2\text{KHCO}_3$

COMMON USES AND POTENTIAL HAZARDS

One of the most familiar applications of potassium bicarbonate is as an antacid to treat the symptoms of upset stomach. The compound reacts with stomach acid—hydrochloric acid;

POTASSIUM BICARBONATE



Potassium bicarbonate. Red atoms are oxygen, white atom is hydrogen; black atom is carbon; and turquoise atom is potassium. Gray stick indicates double bond. PUBLISHERS RESOURCE GROUP

HCl—to relieve gaseous distress, stomach pain, and heartburn. The compound can also be used to treat potassium deficiency in the body. Some research suggests that potassium bicarbonate may help restore muscle and bone tissue, particularly in women with the degenerative bone disease osteoporosis. The compound is also used as a food additive, as a leavening agent, to maintain proper acidity in foods, to supply potassium to a diet, and to provide the bubble and fizz in carbonated drinks.

Potassium bicarbonate is also used in certain types of fire extinguishers. When such an extinguisher is used, the potassium bicarbonate reacts with an acid present in the device to produce carbon dioxide. The carbon dioxide propels a liquid from the extinguisher and, itself, helps put out a fire. Potassium bicarbonate is also used in agriculture to maintain proper acidity in soils and to supply potassium that may be missing from the ground.

Under normal circumstances, potassium bicarbonate poses no health threat to humans. Excess potassium in the body may result in a condition known as hyperkalemia, characterized by tingling of the hands and feet, muscle weakness, and temporary paralysis. Such a condition is very rare when potassium bicarbonate is used in normal amounts.

Interesting Facts

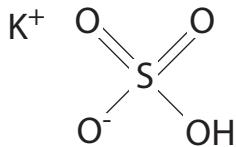
Potassium bicarbonate can be substituted for baking soda (sodium bicarbonate;

NaHCO_3) for people who are on a low sodium diet.

FOR FURTHER INFORMATION

- “Potassium Bicarbonate.” Yale New Haven Health Drug Guide. http://yalenewhavenhealth.org/library/healthguide/en_us/drugguide/topic.asp?hwid=do3600a1 (accessed on October 31, 2005).
- “Potassium Bicarbonate Handbook.” Armand Products Company. Available online at http://www.oxy.com/OXYCHEM/Products/potassium_bicarbonates/literature/PootBiVs6.pdf (accessed on October 31, 2005).
- Rowley, Brian. “Fizz or Sizzle? Potassium Bicarbonate Could Help Spare Muscle and Bone.” *Muscle & Fitness* (December 2002): 72.
- “Strong Muscle and Bones.” *Prevention* (June 1, 1995): 70–73.

See Also Sodium Bicarbonate



OTHER NAMES:	Potassium hydrogen sulfate; potassium acid sulfate
FORMULA:	KHSO_4
ELEMENTS:	Potassium, hydrogen, sulfur, oxygen
COMPOUND TYPE:	Acid salt (inorganic)
STATE:	Solid
MOLECULAR WEIGHT:	136.17 g/mol
MELTING POINT:	about 200°C (about 400°F); decomposes
BOILING POINT:	Not applicable
SOLUBILITY:	Soluble in water; decomposes in alcohol

KEY FACTS

Potassium Bisulfate

OVERVIEW

Potassium bisulfate (poe-TAS-ee-yum BYE-sul-fate) is an odorless white crystalline solid that begins to decompose at its melting point. It is deliquescent, meaning that has such a strong tendency to absorb moisture from the air that it becomes wet and dissolves in the water it has absorbed.

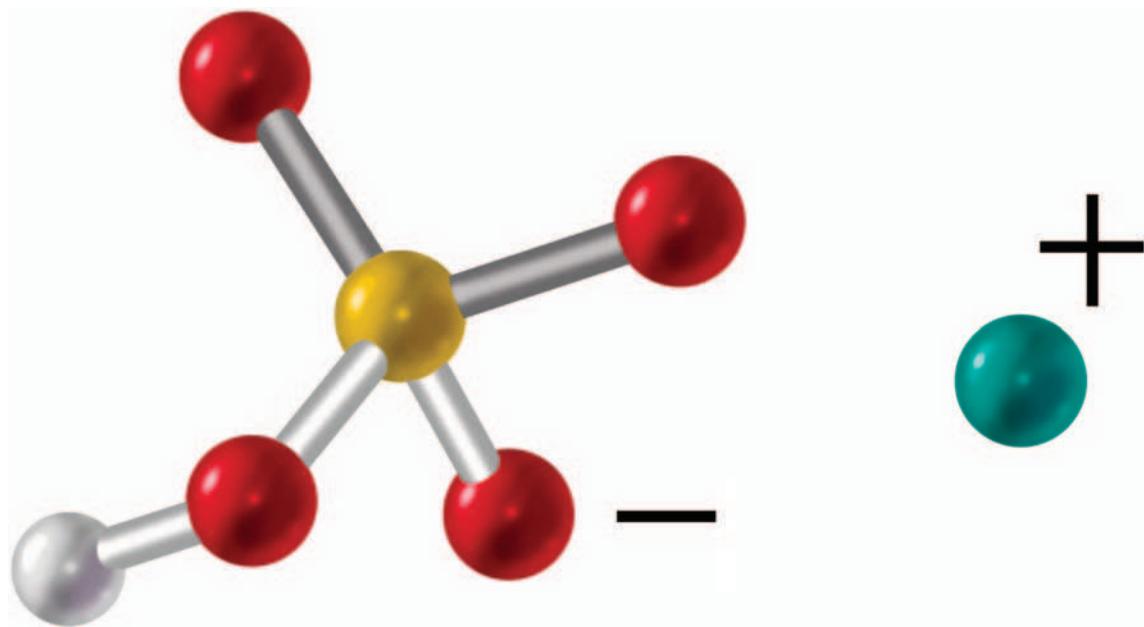
HOW IT IS MADE

Potassium bisulfate is typically made by heating potassium sulfate (K_2SO_4) with sulfuric acid. The acid provides the hydrogen needed to convert the salt (K_2SO_4) to the corresponding acid salt (KHSO_4).

COMMON USES AND POTENTIAL HAZARDS

Potassium bisulfate is used as a food additive. The compound is listed on the U.S. Food and Drug Administration's

POTASSIUM BISULFATE



Potassium bisulfate. Red atoms are oxygen; white atom is hydrogen; yellow atom is sulfur; and turquoise atom is potassium. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

list of Generally Regarded as Safe (GRAS) list. The list contains chemicals thought to be safe for human consumption even though they have not been tested. Potassium bisulfate is used in foods as a preservative because it interferes with the growth of insects, bacteria, and fungi that cause foods to spoil. It is also used as a leavening agent in cake mixes. One of its most important uses is in the wine industry, where it is used to convert certain compounds that occur naturally in grapes into potassium bitartrate. Potassium bisulfate is also used as a flux, in the analysis of ores and silica compounds, in the manufacture of fertilizers, and in the preparation of methyl and ethyl acetate.

Potassium bisulfate is a strong irritant to human tissue. If spilled on the skin, inhaled, or ingested, it can burn tissue causing skin rashes, sore nasal passages, irritation of the throat, and damage to the eyes. Burns of the mouth and stomach may also occur. These hazards are of concern primarily to people who work directly with the compound and do not pose a threat as a food additive.

Words to Know

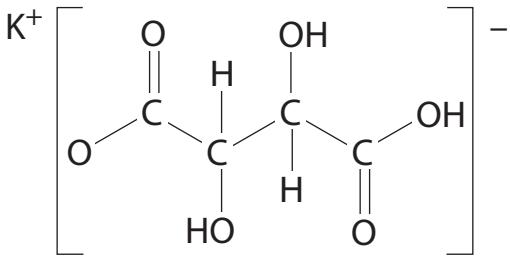
DELIQUESCENT having a strong tendency to absorb moisture from the air, so that it becomes wet and dissolves in the water it has absorbed.

FLUX a material that lowers the melting point of another substance or mixture of substances or that is used in cleaning a metal.

FOR FURTHER INFORMATION

"Agency Response Letter: GRAS Notice No. GRN 000060." U.S. Food and Drug Administration.
http://www.cfsan.fda.gov/~rdb/opa_060.html (accessed on November 1, 2005).

"Safety (MSDS) data for potassium bisulfate ." Physical and Theoretical Chemistry Laboratory, University of Oxford.
http://ptcl.chem.ox.ac.uk/MSDS/PO/potassium_bisulfate.html (accessed on November 1, 2005).



OTHER NAMES:

Potassium hydrogen tartrate; potassium acid tartrate; cream of tartar

FORMULA:
 $\text{KHC}_4\text{H}_4\text{O}_6$

ELEMENTS:

Potassium, hydrogen, carbon, oxygen

COMPOUND TYPE:
Salt (acid salt); inorganic

STATE:
Solid

MOLECULAR WEIGHT:
188.18 g/mol

MELTING POINT:
Not available

BOILING POINT:
Not available

SOLUBILITY:
Somewhat soluble in cold water; very soluble in hot water; insoluble in alcohol

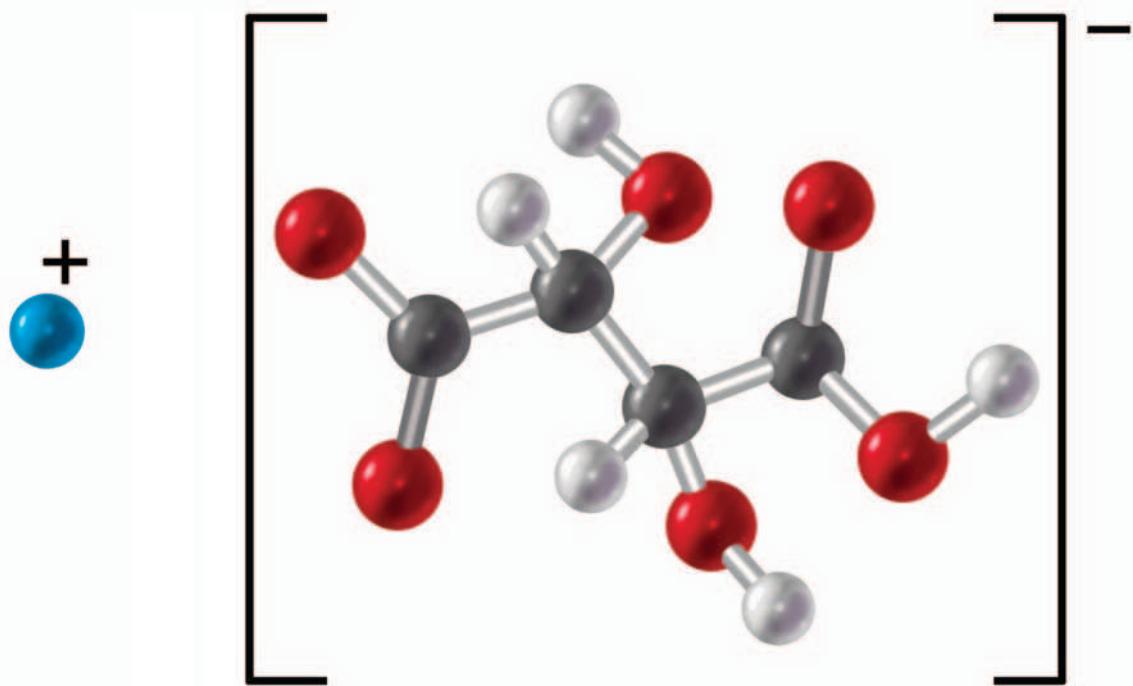
**K
E
Y
F
A
C
T
S**

OVERVIEW

Potassium bitartrate (poe-TAS-ee-yum bye-TAR-trate) is a colorless crystalline or white powdery solid with a pleasant, slightly acidic taste. The compound is a by-product of the fermentation of grape juice and, as such, may have been known to humans for as long as seven thousand years. An article in the journal *Nature* reported some years ago that traces of the calcium salt of tartaric acid, a cousin of potassium bitartrate, was found in remnants of a pottery jar in northern Iran dating to about 7,000 BCE. Potassium bitartrate was used by ancient people in a wide range of household uses, from cooking and baking to cleaning. The true chemical nature of the substance long known as cream of tartar was determined in 1770 by the Swedish chemist Karl Wilhelm Scheele (1742–1786).

HOW IT IS MADE

Potassium bitartrate is made today by the process that has been used for centuries. Wine lees (the solid material left



Potassium bitartrate. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atom is potassium. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

after grapes have been crushed to make wine) are treated with hot water, which dissolves the potassium bitartrate. The hot solution is then allowed to evaporate. As potassium bitartrate crystals form, they are removed and purified.

COMMON USES AND POTENTIAL HAZARDS

The primary use of potassium bitartrate is in commercial food production and household cooking and baking. The compound is added to foods to stabilize egg whites after they have been beaten, to add body to a product, and to produce creamier textures for sugar-based foods. A small amount of potassium bitartrate also adds a pleasantly acidic taste to foods. The compound is found most commonly in baked goods, candies, crackers, confections, gelatins, puddings, jams and jellies, soft drinks, margarines, and frostings. Some of the functions that potassium bitartrate performs in foods include:

- It acts as a leavening agent, causing a product to rise, in baked goods;
- It serves as an anticaking agent and stabilizer to thicken some food products;

- It prevents the crystallization of sugars when making candy and sugary syrups;
- It hides the harsh aftertaste and intensifies the flavors of some foods; and
- It improves the color of vegetables that have been boiled during preparation.

Potassium bitartrate also has a number of other uses. It can be used as a laxative for both humans and domestic animals. It is used in the processing of some metals, giving a colored tinge to the final product. The compound is also an ingredient in products used to clean brass, copper, aluminum, and other metals. And it is used in the chemical industry as a raw material for the preparation of other tartrate compounds.

There are no known health hazards for potassium bitartrate except for the general overall caution given in the Introduction that all chemicals in some concentrations can pose a hazard. Potassium bitartrate is thought to be one of the safest chemical compounds for general use.

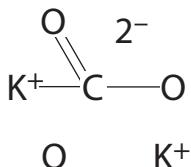
FOR FURTHER INFORMATION

“Potassium Bitartrate.” ChemicalLand21.com.

[http://www.chemicaland21.com/arokorhi/lifescience/foco/
POTASSIUM%20BITARTRATE.htm](http://www.chemicaland21.com/arokorhi/lifescience/foco/POTASSIUM%20BITARTRATE.htm) (accessed on October 29, 2005).

“Potassium Bitartrate.” J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/p5565.htm>
(accessed on October 29, 2005).



OTHER NAMES:	See Overview.
FORMULA:	K_2CO_3
ELEMENTS:	Potassium, carbon, oxygen
COMPOUND TYPE:	Salt (inorganic)
STATE:	Solid
MOLECULAR WEIGHT:	138.21 g/mol
MELTING POINT:	898°C (1650°F)
BOILING POINT:	Not applicable; decomposes above melting point
SOLUBILITY:	Soluble in water; insoluble in ethyl alcohol

KEY FACTS

Potassium Carbonate

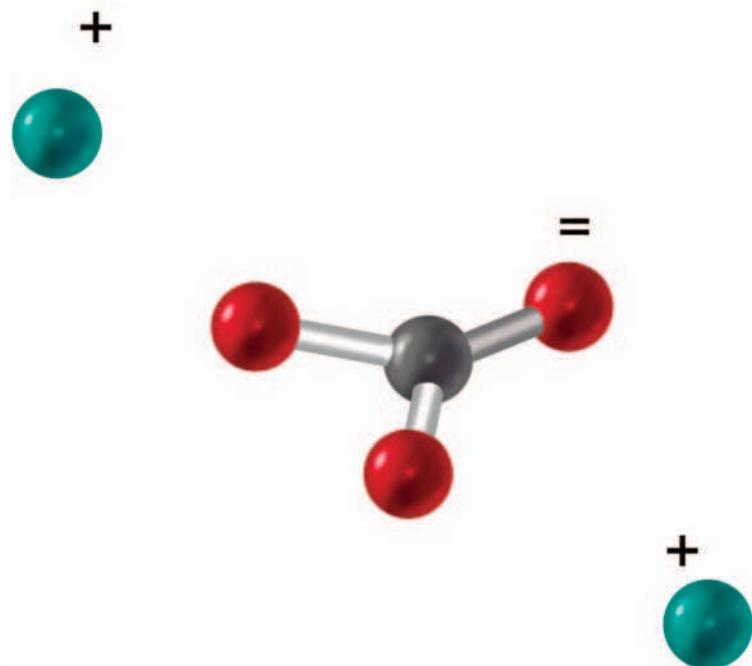
OVERVIEW

Potassium carbonate (poe-TAS-ee-yum KAR-bun-ate) is also known as potash, pearl ash, salt of tartar, carbonate of potash, and salt of wormwood. It is a white, translucent, odorless, granular powder or crystalline material that tends to absorb water from the air. As it does, it is converted into the sesquihydrate (“sesqui” = one-and-a-half) with the formula $\text{K}_2\text{CO}_3 \cdot 1.5\text{H}_2\text{O}$. That formula means that three molecules of potassium carbonate share two molecules of water among them.

Potash is easily produced by pouring water over the ashes of burned plants and then evaporating the solution formed in large pots (hence the name: “pot” “ash”). The process has been known since at least the sixth century CE and the resulting product used in the manufacture of soap. Potash was one of the first chemicals to be exported by American colonists, with shipments having left Jamestown as early as 1608. Potassium carbonate is also called pearl ash and salt of tartar, both of which are impure forms of the compound. The impurities present include sodium chloride

POTASSIUM CARBONATE

Potassium carbonate. Red atoms are oxygen; black atom is carbon; and turquoise atoms are potassium. Gray stick indicates a double bond.
PUBLISHERS RESOURCE GROUP



and some heavy metals (such as iron and lead). The primary uses of potassium carbonate are in the production of fertilizers, soaps, and heat-resistant glass.

HOW IT IS MADE

Most of the potassium carbonate made in the United States is produced beginning with potassium chloride (KCl) obtained from seven mines in New Mexico, Michigan, and Utah. The potassium chloride is first converted to potassium hydroxide (KOH) by electrolysis. The potassium hydroxide is then treated with carbon dioxide (CO_2) to obtain potassium bicarbonate (KHCO_3). Finally, the potassium bicarbonate is decomposed by heating, yielding water, carbon dioxide, and potassium carbonate.

Another method of preparation, called the Engel-Precht process, is a modification of this procedure. A mixture of potassium chloride, magnesium carbonate or magnesium oxide, and carbon dioxide is treated under 30 atmospheres

Interesting Facts

- The first patent ever issued in the United States was awarded in 1790 to Samuel Hopkins for a new and better way of making pearl ash.
- Pearl ash was used in the United States in the eighteenth century as a leavening agent in the baking of bread.
- The demand for potash began to fall off in the late eighteenth century as improved methods for the synthesis of sodium carbonate were developed. Sodium carbonate can replace potassium carbonate in many applications.
- The term potash has historically had many different meanings. It has been used to refer to potassium hydroxide (KOH), potassium chloride (KCl), potassium sulfate (K_2SO_4), potassium nitrate (KNO_3), or to some combination of these compounds.

of pressure, with the formation of a double salt, $KHCO_3 \cdot MgCO_3 \cdot 4H_2O$. The double salt is then heated to obtain potassium carbonate. The traditional method of obtaining potash from wood and vegetable ash is now obsolete.

COMMON USES AND POTENTIAL HAZARDS

An estimated 7 million metric tons (6.5 million short tons) of potassium carbonate were produced in the United States in 2005. Of that amount, nearly 90 percent was used for the production of fertilizers. Potash provides plants with the potassium they need to stay healthy and grow. Potassium is one of the three major nutrients required by plants, the other two being nitrogen and phosphorus. The next largest application for potassium carbonate is in the chemical industry, where it is used as a raw material to make other chemical compounds, potassium silicate being the most common.

Smaller amounts of potassium carbonate are still used for what was once its major application: the manufacture of soap. Potassium soaps (made from potassium carbonate) have

Words to Know

CAUSTIC Strongly basic or alkaline; can irritate or corrode living tissue.

ELECTROLYSIS Process in which an electric current is used to bring about chemical changes.

FLUX A material that lowers the melting point of another substance or mixture of substances or that is used in cleaning a metal.

some characteristics different from more common sodium soaps (made from sodium carbonate). They tend to be softer or even liquid and better able to create suds in water that contains a high concentration of minerals. Potassium carbonate is also used to make specialty glasses, such as television screens, cathode ray tubes, and optical lenses. Some other uses of the compound include:

- For glazes in the making of pottery;
- In the manufacture of pigments and printing inks;
- As an additive in certain food products, chocolate being one example;
- For the tanning and finishing of leather and the dyeing, washing, and finishing of wool; and
- As a flux in metal working.

Potassium carbonate in dry or dissolved form is an irritant to the eyes, skin, and respiratory system. It can cause inflammation of the skin, eyes, throat, and stomach. Potassium carbonate's action is caused by its caustic properties in water solution, which are produced when it is dissolved in water or when it is absorbed by moist tissues in the body.

FOR FURTHER INFORMATION

“International Potash Institute.”

<http://www.ipipotash.org/index.php> (accessed on October 31, 2005).

“The Potash Trade.” Townships Heritage WebMagazine.

<http://www.townshipheritage.com/Eng/Hist/Life/potash.html> (accessed on October 31, 2005).

“Potassium Carbonate.” Chemical Land 21.

<http://www.chemicalland21.com/arokorhi/industrialchem/inorganic/POTASSIUM%20CARBONATE.htm>(accessed on October 31, 2005).

“Potassium Carbonate.” J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/p5609.htm>
(accessed on October 31, 2005).

Willett, Jason C. “Potash.” U.S. Geological Survey Commodity Statistics and Information. Available online at <http://minerals.usgs.gov/minerals/pubs/commodity/potash/potasmybo4.pdf> (accessed on October 31, 2005).

See Also Potassium Chloride; Potassium Hydroxide



OTHER NAMES:	Potassium muriate; muriate of potash
FORMULA:	KCl
ELEMENTS:	Potassium, chlorine
COMPOUND TYPE:	Binary salt (inorganic)
STATE:	Solid
MOLECULAR WEIGHT:	74.55 g/mol
MELTING POINT:	771°C (1420°F)
BOILING POINT:	Not applicable; sublimes at about 1500°C (2700°F)
SOLUBILITY:	Very soluble in water; slightly soluble in ethyl alcohol, and insoluble in ether, acetone, and other organic solvents

KEY FACTS

Potassium Chloride

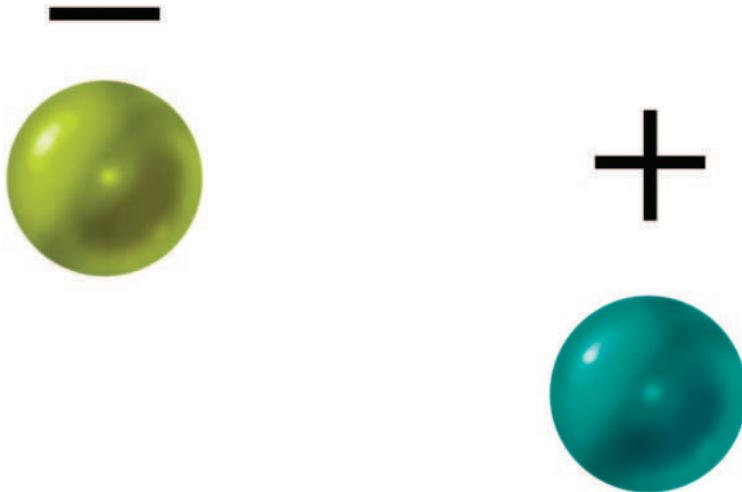
OVERVIEW

Potassium chloride (poe-TAS-ee-yum KLOR-ide) occurs as a white or colorless crystalline solid or powder. It is odorless, but has a strong saline (salty) taste. It occurs naturally in the minerals sylvite, carnallite, kainite, and sylvinitite. It also occurs in sea water at a concentration of about 0.076 percent (grams per milliliter of solution). Potassium chloride is the most abundant compound of the element potassium and has the greatest number of applications of any salt of potassium. By far the most important application of potassium chloride is in the manufacture of fertilizers.

HOW IT IS MADE

All of the major sources of potassium chloride have their origin in sea water. Sea water is a solution of a number of salts dissolved in water. The most important of those salts are sodium chloride (about 2.3 percent), magnesium chloride (about 0.5 percent), sodium sulfate (about 0.4 percent),

Potassium chloride. Turquoise atom is potassium and green atom is chlorine. Potassium atom is positively charged. Chlorine atom is negatively charged. PUBLISHERS RESOURCE GROUP



calcium chloride (about 0.1 percent) and potassium chloride (about 0.07 percent). When large bodies of sea water dry up, they leave behind complex mixtures of minerals consisting of these salts. Over millions of years, huge deposits of these minerals have been buried under the land. In the United States, sea salt deposits are found in New Mexico, Texas, California, and Michigan.

Any one of the salts present in a sea salt deposit—including potassium chloride—can be extracted by a common procedure. The minerals that make up the deposit are crushed and dissolved in hot water. The solution is then allowed to cool very slowly. As it cools, each of the dissolved salts crystallizes out at a specific temperature, is removed from the solution, and is purified. Since potassium chloride is much more soluble in hot water than in cold water, it crystallizes out after other salts have been removed.

The majority of potassium chloride in the United States is now extracted by a lengthy process that also begins with the crushing of natural ores, such as sylvite and carnalite. The solid mixture is then cleaned and purified before being treated with a flotation agent, usually some type of amine. A flotation agent is a material that coats the desired compound, such as potassium chloride, and allows it to float to the surface of the reaction chamber, like the soap suds that float on top of a washing machine. An amine is an organic compound that contains the nitrogen, usually as the -NH_2 ,

Interesting Facts

- One use of potassium chloride is as a lethal injection for prisoners who have been given the death penalty. The chemical interferes with normal heart function and causes a heart attack within five

to about eighteen minutes after injection. Thirty-four of the United States prescribe death by lethal injection for prisoners who have been convicted of murder.

group. The amine-coated potassium chloride is skimmed off the top of the reaction mixture, purified, and prepared in some crystalline or powder form.

COMMON USES AND POTENTIAL HAZARDS

Potassium chloride is present in some foods in small amounts. The compound is also used as a food additive to increase the acidity and to stabilize, thicken, or soften some food products, such as jams and jellies and preserves that are artificially sweetened. Many infant formulas also contain potassium chloride. Potassium chloride is also used as a nutrient for yeast cultures and in making beer. The compound is used as a salt substitute for people who are on low-salt (meaning low-sodium) diets. Some brand names of these products are LoSalt®, Reheis Less Salt Blend®, and Morton® Lite Salt®.

The largest application of potassium chloride is in the production of fertilizers. More than ninety percent of the potassium chloride produced in the United States is used for that purpose. The compound provides the potassium plants need to stay healthy and grow normally. It is one of three macronutrients—substances needed in relatively large amounts—for normal growth. The other two macronutrients are phosphorus and nitrogen. Smaller amounts of potassium chloride are used in the production of other potassium compounds, in photography, and in chemical research applications.

FOR FURTHER INFORMATION

Benfell, Carol. "Routine but Deadly Drug: Potassium Chloride Has a Jekyll and Hyde Personality." American Iatrogenic Association.
<http://www.iatrogenic.org/potchlor.html> (accessed on October 31, 2005).

"Potassium Chloride." MedicineNet.com.
http://www.medicinenet.com/potassium_chloride/article.htm (accessed on October 31, 2005).

"Potassium Chloride." University of Maryland Medical Center.
<http://www.umm.edu/altmed/ConsDrugs/PotassiumChloridecd.html> (accessed on October 31, 2005).

See Also Potassium Carbonate; Potassium Hydroxide;
Potassium Nitrate; Sodium Chloride

OTHER NAMES:	Potassium monofluoride
FORMULA:	KF
ELEMENTS:	Potassium, fluorine
COMPOUND TYPE:	Binary salt (inorganic)
STATE:	Solid
MOLECULAR WEIGHT:	58.10 g/mol
MELTING POINT:	858°C(1580°F)
BOILING POINT:	1502°C(2736°F)
SOLUBILITY:	Soluble in cold water; very soluble in hot water; insoluble in ethyl alcohol; soluble in hydrofluoric acid (H ₂ F ₂)

KEY FACTS

Potassium Fluoride

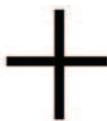
OVERVIEW

Potassium fluoride (poe-TAS-ee-yum FLU-ride) is a colorless or white crystalline or powdery compound with no odor, but a sharp, salty taste. It has somewhat limited uses in industry and chemical research.

HOW IT IS MADE

In one method for making potassium fluoride, potassium carbonate (K₂CO₃) is dissolved in hydrofluoric acid, resulting in the formation of potassium bifluoride (KHF₂): K₂CO₃ + 2H₂F₂ → 2KHF₂ + CO₂ + H₂O. The potassium bifluoride is then heated to form potassium fluoride and hydrogen fluoride: KHF₂ → KF + HF.

Potassium fluoride can also be prepared by the direct reaction between hydrofluoric acid and potassium hydroxide: H₂F₂ + 2KOH → 2KF + 2H₂O. The potassium fluoride thus formed is then dried and crystallized or converted to powder form.



Potassium fluoride. Silver atom is potassium; yellow atom is fluorine. PUBLISHERS RESOURCE GROUP



COMMON USES AND POTENTIAL HAZARDS

Potassium fluoride is used as a fluoridating agent—a substance that provides fluorine atoms to other compounds—in the preparation of organic chemicals. It also finds some use in the field of metallurgy, where it is used as a flux, to finish metals, to make coatings for metals, and in tin

Interesting Facts

- In France, potassium fluoride is sometimes added to table salt to help prevent dental cavities.
- Potassium fluoride can not be shipped out of the United States to other

countries without a special license from the U.S. Department of Commerce because the compound is an important raw material in the manufacture of certain chemical weapons.

Words to Know

FLUX A material that lowers the melting point of another substance or mixture of substances or that is used in cleaning a metal.

METALLURGY The study of the properties and structures of metals.

plating. Potassium fluoride is used to frost and etch glass, as in the manufacture of some optical glasses, and to make insecticides, pesticides, and disinfectants.

Potassium fluoride is irritating to the skin, eyes, and respiratory system. It is moderately toxic by ingestion, causing nausea, vomiting, diarrhea, and stomach pains. In larger doses, it can cause damage to the brain, kidneys, and heart. Long-term exposure to potassium fluoride can cause damage to the teeth and bones. One condition that can develop is called fluorosis. Symptoms of the condition include brittle bones, weight loss, anemia, hardening of the ligaments, and stiffness of the joints.

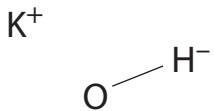
FOR FURTHER INFORMATION

Hernandez, Lucula Pazos. "Prevention of Dental Caries Through Salt Fluoridation in Mexico." <http://www.ibiblio.org/taft/cedros/english/newsletter/n3/prevent.html> (accessed on October 31, 2005).

"Potassium Fluoride." Patnaik, Pradyot. *Handbook of Inorganic Chemicals*. New York: McGraw Hill, 2003, 754-755.

"Potassium Fluoride, Anhydrous." J. T. Baker. <http://www.jtbaker.com/msds/englishhtml/p5774.htm> (accessed on October 31, 2005).

"Silver Production on the Moon." The Artemis Project. http://asi.org/adb/02/13/02/silicon_production.html (accessed on October 31, 2005).



OTHER NAMES:
Caustic potash;
potash lye; potassa;
potassium hydrate

FORMULA:
 KOH

ELEMENTS:
Potassium, oxygen,
hydrogen

COMPOUND TYPE:
Base (inorganic)

STATE:
Solid

MOLECULAR WEIGHT:
56.10 g/mol

MELTING POINT:
406°C (763°F)

BOILING POINT:
1327°C (2421°F)

SOLUBILITY:
Soluble in water,
ethyl alcohol, methyl
alcohol, and glycerol

K E Y F A C T S

OVERVIEW

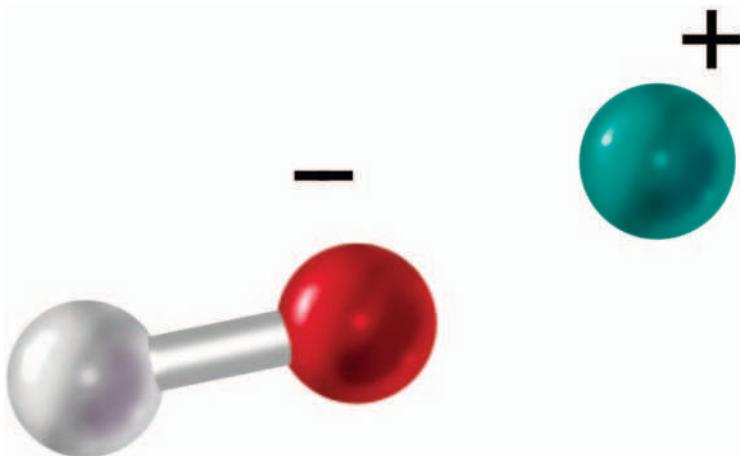
Potassium hydroxide (poe-TAS-ee-yum hy-DROK-side) is a white deliquescent solid that is available in sticks, lumps, flakes, or pellets. A deliquescent material is one that tends to absorb so much moisture from the atmosphere that it becomes very wet, even to the point of dissolving in the water it has absorbed. Potassium hydroxide also absorbs carbon dioxide from the air, changing in the process to potassium carbonate (K_2CO_3). Potassium hydroxide is one of the most caustic materials known. It has a number of uses in industry and agriculture.

Potassium hydroxide is chemically very active. It reacts violently with acids, generating significant amounts of heat in the process. In moist air, it corrodes metals such as tin, lead, zinc, and aluminum with the release of combustible and explosive hydrogen gas.

HOW IT IS MADE

Potassium hydroxide is made by the electrolysis of an aqueous solution of potassium chloride (KCl). In that process,

Potassium hydroxide. Red atom is oxygen; white atom is hydrogen; and turquoise atom is potassium. PUBLISHERS RESOURCE GROUP



an electric current decomposes potassium chloride into potassium and chlorine. The chlorine escapes as a gaseous by-product and the potassium reacts with water to form potassium hydroxide.

COMMON USES AND POTENTIAL HAZARDS

An estimate 440,000 metric tons (485,000 short tons) of potassium hydroxide were used in the United States in 2005. About 53 percent of that amount was used in the production of other potassium compounds, especially potassium carbonate (28 percent), potassium acetate, potassium cyanide, potassium permanganate, and potassium citrate. About 10 percent of all caustic potash was used in the manufacture of potassium soaps and detergents. Most soaps and detergents are made of sodium hydroxide. But potassium hydroxide can be substituted for sodium hydroxide to obtain soaps and detergents with special properties. Liquid soaps and soaps that will lather in salt water or water with a high mineral content are examples of such specialized potassium soaps.

Some other applications of potassium hydroxide include:

- The manufacture of liquid fertilizers, herbicides, and other agricultural chemicals;
- As a neutralizing agent in many chemical and industrial processes;
- In the production of synthetic rubber;

Interesting Facts

- Pure potassium hydroxide is difficult to prepare since the compound is so reactive that it tends to react with moisture, carbon dioxide, and other impurities with which it comes into contact. The compound is commercially available in a purity of about 90 percent. Much purer products are available, however, when needed.
- In the food production industry, where it is used in the removal of peels from fruits and vegetables, the caramelization of products that contain sugar, the thickening of ice cream, the softening of olives, the production of chocolate and cocoa, and the preparation of hominy from corn kernels;
- In the manufacture of alkaline storage batteries and some types of fuel cells;
- In the refining of petroleum;
- As a way for removing horn buds from young cattle;
- In a number of cosmetic procedures, such as softening of cuticles, removal of warts, and cleaning of dentures; and
- As an ingredient in paint removers.

Potassium hydroxide is a very hazardous chemical. It is corrosive to tissue and can cause severe burns of the skin, eyes, and mucous membranes. If ingested, it can cause internal bleeding, scarring of tissue, nausea, vomiting, diarrhea, and lowered blood pressure that can result in a person's collapse. In sufficient amounts, it can cause death. Inhalation of potassium hydroxide fumes or dust can cause lung irritation, sneezing, sore throat, runny nose, and severe damage to the lungs. In contact with the eyes, the compound can cause blurred vision and, in sufficient amounts, loss of eyesight. People who have to work with the compound should always wear goggles, gloves, and protective clothing to reduce their risk of contact with the chemical.

Words to Know

AQUEOUS SOLUTION A solution that consists of some material dissolved in water.

CAUSTIC Strongly basic or alkaline; able to irritate or corrode living tissue.

DELIQUESCENT Having the tendency to absorb moisture and, therefore, dissolve or melt.

MUCOUS MEMBRANES Tissues that line the moist inner lining of the digestive, respiratory, urinary and reproductive systems.

FOR FURTHER INFORMATION

“Caustic Potash.” Occidental Petroleum Corporation.

http://www.oxy.com/OXYCHEM/Products/caustic_potash/caustic_potash.htm (accessed on November 1, 2005).

Cavitch, Susan Miller. *The Natural Soap Book: Making Herbal and Vegetable Based Soaps*. Markham, Canada: Storey Publishing, 1995.

“Potassium Hydroxide.” Medline Plus.

<http://www.nlm.nih.gov/medlineplus/ency/article/oo2482.htm> (accessed on November 1, 2005).

“Potassium Hydroxide.” NIOSH Pocket Guide to Chemical Hazards.

<http://www.cdc.gov/niosh/npg/npgd0523.html> (accessed on November 1, 2005).

See Also Sodium Hydroxide

FORMULA:

KI

ELEMENTS:

Potassium, iodine

COMPOUND TYPE:
Binary salt (inorganic)**STATE:**
Solid**MOLECULAR WEIGHT:**
166.00 g/mol**MELTING POINT:**
681°C (1260°F)**BOILING POINT:**
1323°C (2413°F)**SOLUBILITY:**
Soluble in water,
ethyl alcohol, acetone, and glycerol**K E Y F A C T S**

Potassium Iodide

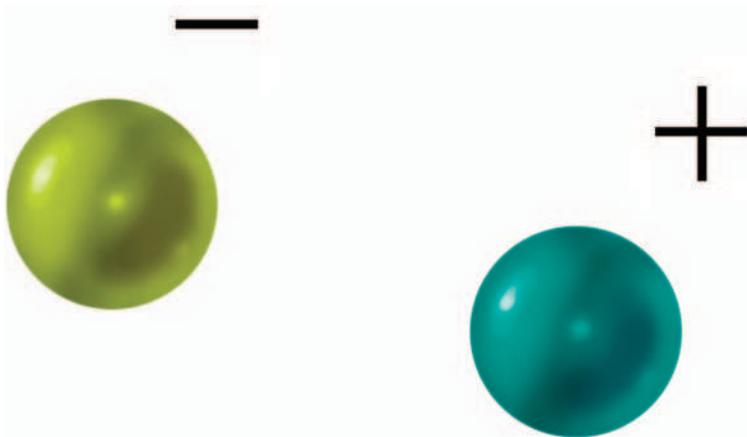
OVERVIEW

Potassium iodide (poe-TAS-ee-yum EYE-oh-dide) is a white crystalline, granular, or powdered solid with a strong, bitter, salty taste. It is used as a feed additive, a dietary supplement, in photographic films, and in chemical research.

HOW IT IS MADE

A number of methods are available for the preparation of potassium iodide. In one procedure, elemental iodine (I_2) is added to a solution of potassium hydroxide (KOH): $I_2 + 6KOH \rightarrow 5KI + KIO_3 + 3H_2O$. The potassium iodide formed is separated from the potassium iodate (KIO_3) by fractional crystallization. That is, the solution is warmed and then cooled. As the temperature falls, the two compounds, potassium iodide and potassium iodate, crystalize out at different temperatures and can be separated from each other. The potassium iodate can then be heated, causing it to decompose and make additional potassium iodide: $2KIO_3 \rightarrow 2KI + 3O_2$.

Potassium iodide. Green atom is iodine and turquoise atom is potassium. PUBLISHERS RESOURCE GROUP



Potassium iodide can also be produced by reacting hydriodic acid (HI) with potassium bicarbonate: $\text{HI} + \text{KHCO}_3 \rightarrow \text{KI} + \text{CO}_2 + \text{H}_2\text{O}$.

Finally, the compound can be made by reacting iron (III) iodide with potassium carbonate (K_2CO_3): $\text{Fe}_3\text{I}_8 + 4\text{K}_2\text{CO}_3 \rightarrow 8\text{KI} + 4\text{CO}_2 + \text{Fe}_3\text{O}_4$.

COMMON USES AND POTENTIAL HAZARDS

Potassium iodide is added to animal feeds to ensure that domestic animals get the iodine they need in their daily diets. A mixture of iodine in aqueous potassium iodide called SSKI has long been used as a disinfectant. The active agent in this mixture is iodine, and the potassium iodide is added to increase the iodine's solubility in water. SSKI is used to purify small amounts of water, to clear up pimples, to prevent sinus infections, and to treat bladder, lung, and stomach infections.

Perhaps the best known use of potassium iodide today is as a treatment for radiation exposure. When a nuclear bomb explodes or a nuclear accident occurs, one of the most dangerous products released to the environment is a radioactive isotope known as iodine-131. Iodine-131 enters the human body and travels to the thyroid, where it attacks cells and tissues, eventually resulting in thyroid cancer. Experts recommend that people exposed to radiation take potassium iodide as a protection against this hazard. The potassium iodide saturates

Interesting Facts

- Iodized salt contains a small amount of potassium iodide. The potassium iodide is added to ensure that people get enough iodine in their daily diet. A deficiency of iodine results in a medical condition known as goiter, a large swelling in the neck.

Before iodized salt was available, many people failed to get enough iodine in their diets and goiter was a very common health problem.

- Potassium iodide occurs naturally in seaweed.

the thyroid gland, making it difficult for the gland to absorb radioactive iodine that may enter the body.

Potassium iodide is also used in the production of photographic film and in a number of chemical tests, such as the determination of dissolved oxygen in water and the presence of starch in an unknown mixture.

Under most circumstances, there are no health hazards associated with potassium iodide. Taking an excess of the compound may have harmful effects on the thyroid gland, however. For that reason, people with an overactive thyroid should not take potassium iodide unless so directed by their doctors. Also, a person should not take potassium iodide as a preventative treatment against radiation. It provides no protection in advance of radiation exposure and, in excessive amounts, can create problems of its own for the thyroid.

Words to Know

RADIOACTIVE ISOTOPE A form of an element that gives off some form of radiation and changes into another element.

FOR FURTHER INFORMATION

“Frequently Asked Questions on Potassium Iodide (KI).” U.S. Food and Drug Administration.

http://www.fda.gov/cder/drugprepare/KI_Q&A.htm (accessed on November 1, 2005).

“Potassium Iodide.” J. T. Baker.

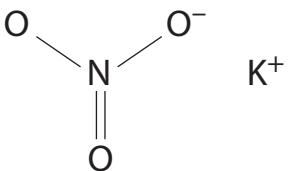
<http://www.jtbaker.com/msds/englishhtml/p5906.htm> (accessed on November 1, 2005).

“Potassium Iodide (Systemic).” Medline Plus.

<http://www.nlm.nih.gov/medlineplus/druginfo/uspdi/202472.html> (accessed on November 1, 2005).

Wright, Jonathan V. “One Mineral Can Help a Myriad of Conditions from Atherosclerosis to ‘COPD’ to Zits.” Tahoma Clinic. <http://www.tahoma clinic.com/iodide.shtml> (accessed on November 1, 2005).

See Also Silver Iodide

**OTHER NAMES:**

Niter; saltpeter;
nitrate of potash

FORMULA:

KNO_3

ELEMENTS:

Potassium, nitrogen,
oxygen

COMPOUND TYPE:

Salt (inorganic)

STATE:

Solid

MOLECULAR WEIGHT:

101.10 g/mol

MELTING POINT:

337°C (639°F)

BOILING POINT:

Not applicable;
decomposes above
400°C (750°F)

SOLUBILITY:

Soluble in water
and glycerol;
slightly soluble in
ethyl alcohol

KEY FACTS

Potassium Nitrate

OVERVIEW

Potassium nitrate (poe-TAS-ee-yum NYE-trate) is transparent, colorless, or white, and may be crystalline or powdery solid. It is odorless with a sharp, cool, salty taste. It is slightly hygroscopic, that is, having a tendency to absorb moisture from the air. Potassium nitrate, more commonly known as saltpeter or niter, has been used by humans for many centuries. Going back as far as ancient Chinese civilizations, the compound was used as an ingredient in fireworks, to preserve foods, to make incense burn more evenly, to increase the male sex drive, and for magic potions.

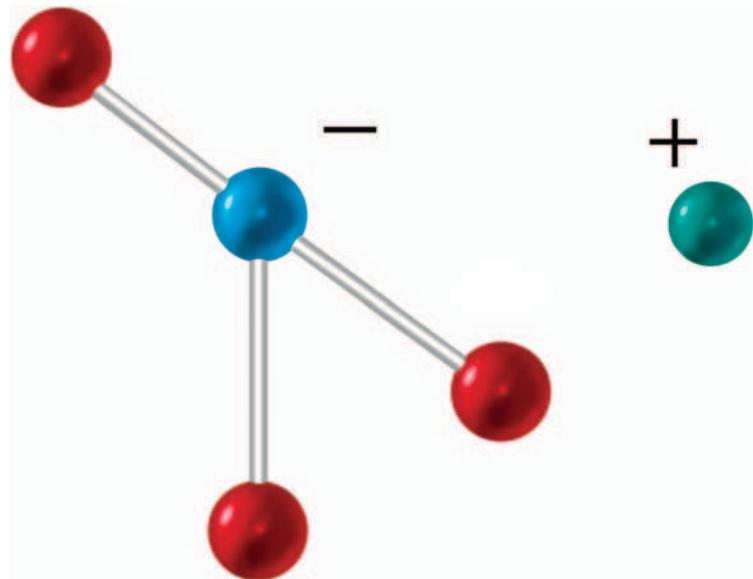
HOW IT IS MADE

Potassium nitrate is made commercially by reacting potassium chloride (KCl) with nitric acid (HNO_3) at high temperatures: $3\text{KCl} + 4\text{HNO}_3 \rightarrow 3\text{KNO}_3 + \text{Cl}_2 + \text{NOCl} + 2\text{H}_2\text{O}$.

The compound can also be obtained for use from natural sources. It occurs as a thin, whitish, glassy crust on rocks in

POTASSIUM NITRATE

Potassium nitrate. Red atoms are oxygen; blue atom is nitrogen; and turquoise atom is potassium. PUBLISHERS RESOURCE GROUP



sheltered areas, such as caves. In warm climates, potassium nitrate forms when bacteria decompose animal feces and other organic matter. The compound usually appears as a white powder on the surface of soil. These sources of potassium nitrate are of use on a small scale basis and have no commercial value.

COMMON USES AND POTENTIAL HAZARDS

The primary use of potassium nitrate is in explosives, blasting powders, gunpowder, fireworks, and matches. The compound is used as an oxidizing agent in such preparations. An oxidizing agent is a substance that provides oxygen for the combustion of some other material. For example, gunpowder, the oldest known explosive, is a mixture of potassium nitrate, charcoal (nearly pure carbon), and sulfur. When the mixture is ignited, the carbon and sulfur burn very rapidly to produce carbon dioxide (CO) and sulfur dioxide (SO_2). At the same time, the potassium nitrate decomposes to produce a variety of products, one of which is nitric oxide (NO). The rapid formation of very hot gases is responsible for the shock wave produced in the explosion.

Interesting Facts

- At one time, potassium nitrate was prepared by mixing manure with mortar or wood ash, soil, and an organic material, such as straw. The bed was kept moist with urine and turned often to speed decomposition of the organic matter. After a year, the bed was thoroughly watered, dissolving the potassium nitrate that had accumulated. It was then recrystallized and purified.
- In 1862, leaders of the Confederate Army ordered a chemistry professor at South Carolina College to teach farmers how to make potassium nitrate to ensure an adequate supply of the compound for use in making gunpowder.
- Gunpowder was probably first used as early as the eleventh century. The English natural philosopher Roger Bacon (1214-1294) described a method for making gunpowder in 1242.

Some other uses of potassium nitrate include:

- As a meat preservative that helps meats retain their bright red color;
- As a flux for soldering;
- In fertilizers, especially for use with crops such as tomatoes, potatoes, tobacco, leafy vegetables, citrus fruits, and peaches;
- In the manufacture of glasses and ceramics;
- As an additive for tobacco products that helps the tobacco burn more cleanly and smoothly;
- As an oxidizing agent in rocket propulsion systems;
- As a diuretic, a substance that increases the flow of urine from the body; and
- As a raw material in the manufacture of other potassium compounds.

Exposure to moderate amounts of potassium nitrate dust and fumes can result in irritation of the skin, eyes, and respiratory system. Symptoms may include sneezing, coughing,

Words to Know

FLUX A material that lowers the melting point of another substance or mixture of

substances or that is used in cleaning a metal.

dizziness, drowsiness, and headache. Ingestion of the compound may result in nausea, vomiting, and severe abdominal pain. Exposure to large quantities of the compound may have more serious consequences because it interferes with the blood's ability to transport oxygen. In such cases, a person may experience shortness of breath, bluish skin, serious damage to the kidneys, unconsciousness, and even death. People who work directly with potassium nitrate are at greatest risk for such health problems.

FOR FURTHER INFORMATION

"Material Safety Data Sheet: Potassium Nitrate." Department of Chemistry, Iowa State University.

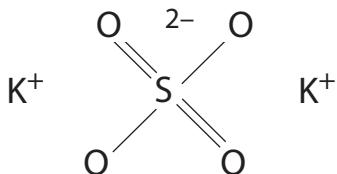
<http://avogadro.chem.iastate.edu/MSDS/KNO3.htm> (accessed on November 3, 2005).

Multhauf, Robert P., and Christine M. Roane. "Nitrates." In *Dictionary of American History*. Edited by Stanley I. Kutler. 3rd ed., vol. 6. New York: Charles Scribner's Sons, 2003.

"Potassium Nitrate." Hazardous Substances Data Bank.
<http://toxnet.nlm.nih.gov/cgi-bin/sis/search/r?dbs+hsdb:@term+@na+Potassium+Nitrate> (accessed on November 3, 2005).

"Potassium Nitrate." International Chemical Safety Cards.
<http://www.inchem.org/documents/icsc/icsc/eicso184.htm> (accessed on November 3, 2005).

"Potassium Nitrate." Skylighter.
http://www.skylighter.com/potassium_nitrate.html (accessed on November 3, 2005).



OTHER NAMES:
See Overview.

FORMULA:
 K_2SO_4

ELEMENTS:
Potassium, sulfur,
oxygen

COMPOUND TYPE:
Salt (inorganic)

STATE:
Solid

MOLECULAR WEIGHT:
174.26 g/mol

MELTING POINT:
1069°C (1956°F)

BOILING POINT:
Vaporizes at 1689°C
(3072°F)

SOLUBILITY:
Soluble in water;
slightly soluble in
glycerol; insoluble in
ethyl alcohol, acetone,
and most other
organic solvents

KEY FACTS

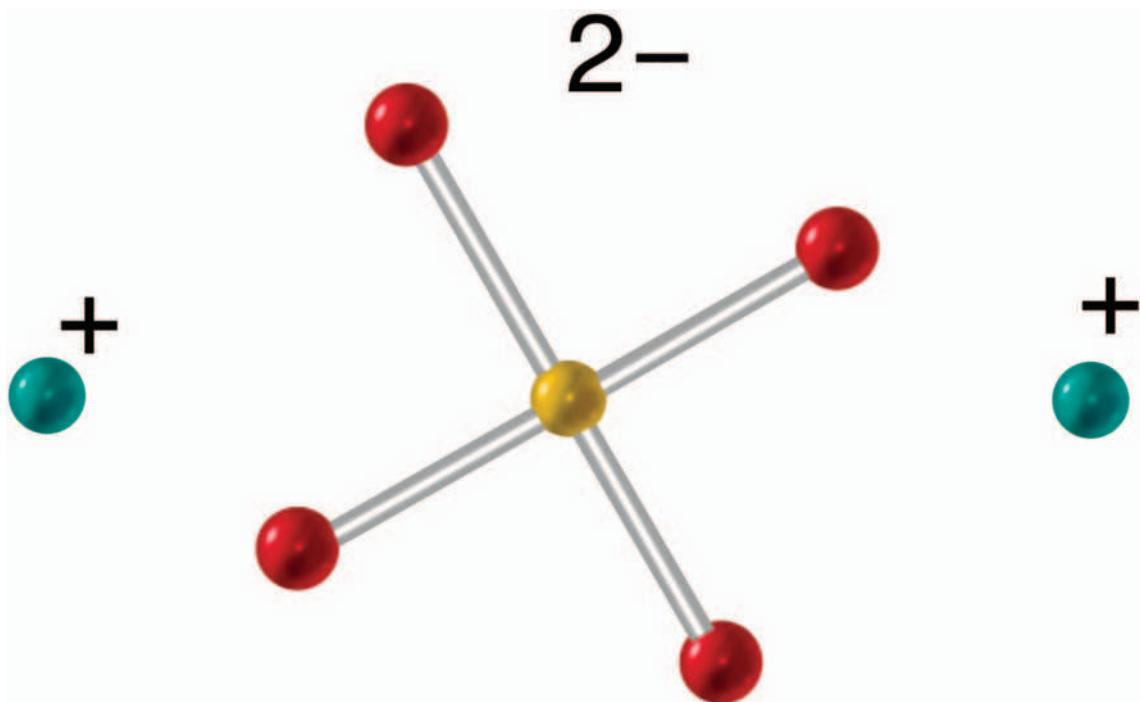
Potassium Sulfate

OVERVIEW

Potassium sulfate (poe-TAS-ee-yum SUL-fate) is also known as potash of sulfur, sulfuric acid dipotassium salt, arcanum duplicatum, and sal polychrestum. It is a colorless or white granular, crystalline, or powdery solid with a bitter, salty taste. It occurs in nature as the mineral arcanite and in the mineral langbeinitite ($K_2Mg_2(SO_4)_3$). The compound was known to alchemists as early as the fourteenth century, and was analyzed by a number of early chemists, including Johann Glauber (1604–1670), Robert Boyle (1627–1691), and Otto Tachenius (c. 1620–1690).

HOW IT IS MADE

A variety of methods for preparing potassium sulfate is available. In one process, the compound is extracted from the mineral langbeinitite by crushing and washing the mineral and then separating out the double salt, $K_2Mg_2(SO_4)_3$. The product is then treated with an aqueous solution of potassium



Potassium sulfate. Red atoms are oxygen; yellow atom is sulfur; and turquoise atoms are potassium. PUBLISHERS RESOURCE GROUP

chloride (KCl) to separate the two parts of the double salt from each other: $K_2Mg_2(SO_4)_3 + 4KCl \rightarrow 3K_2SO_4 + 2MgCl_2$. The compound can also be produced synthetically by treating potassium chloride with sulfuric acid (H_2SO_4): $2KCl + H_2SO_4 \rightarrow K_2SO_4 + 2HCl$.

In a variation of this procedure, potassium chloride is treated with the raw materials from which sulfuric acid is made, rather than the acid itself: $4KCl + 2SO_2 + 2H_2O + O_2 \rightarrow 2K_2SO_4 + 4HCl$.

COMMON USES AND POTENTIAL HAZARDS

Over 90 percent of the potassium sulfate produced in the United States is used as a fertilizer. It provides plants with two essential elements: potassium and sulfur. It finds its greatest use on crops that are sensitive to the chloride ion (Cl^-) present in most conventional agricultural fertilizers. Those crops include coffee, tea, tobacco, citrus fruits, grapes, and potatoes. However, its use is somewhat limited because it is twice as expensive as fertilizers that contain potassium chloride.

Interesting Facts

Potassium sulfate, sodium sulfate, and their double salts with calcium and magnesium occur naturally in salt lakes and in volcanic

lava. At one time and on some uncommon occasions today potassium sulfate was obtained from salt lake sources.

The second most important use of potassium sulfate is as a supplement for animal feeds, accounting for another 8 percent of the compound produced in the United States. The remaining 1 percent of potassium sulfate goes to the production of gypsum board and gypsum cement, for the synthesis of potassium alum (potassium aluminum sulfate), in the manufacturing of glass and ceramics, for the production of dyes and lubricants, and as a flash suppressant in explosives. A flash suppressant is, as its name suggests, a chemical that reduces the amount of flash produced when an explosive is detonated.

Exposure to moderate amounts of potassium sulfate appears to have no serious effects on human health. Ingestion of large amounts of the compound, can cause severe gastrointestinal irritation that requires medical attention.

Words to Know

AQUEOUS SOLUTION A solution that consists of some material dissolved in water.

from simple beginning chemicals, or reactants.

SYNTHESIS A chemical reaction in which some desired chemical product is made

FOR FURTHER INFORMATION

“Potassium Sulfate.” Hummel Croton, Inc.

http://www.hummelcroton.com/msds/k2so4_m.html (accessed on November 3, 2005).

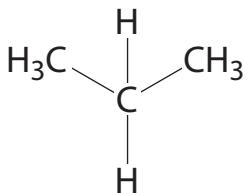
“Potassium Sulfate.” International Programme on Chemical Safety.

<http://www.inchem.org/documents/icsc/icsc/eics1451.htm> (accessed on November 3, 2005).

“Potassium Sulfate.” J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/p6137.htm> (accessed on November 3, 2005).

See Also Potassium Carbonate, Potassium Chloride, Potassium Nitrate

**FORMULA:** C_3H_8 **ELEMENTS:**

Carbon, hydrogen

COMPOUND TYPE:

Alkane; saturated hydrocarbon; organic

STATE:

Gas

MOLECULAR WEIGHT:

44.10 g/mol

MELTING POINT: -187.63°C
 $(-323.73^{\circ}\text{F})$ **BOILING POINT:**
 -42.1°C (-43.8°F)**SOLUBILITY:**
Slightly soluble in water; soluble in ethyl alcohol, ether, benzene**K E Y F A C T S**

Propane

OVERVIEW

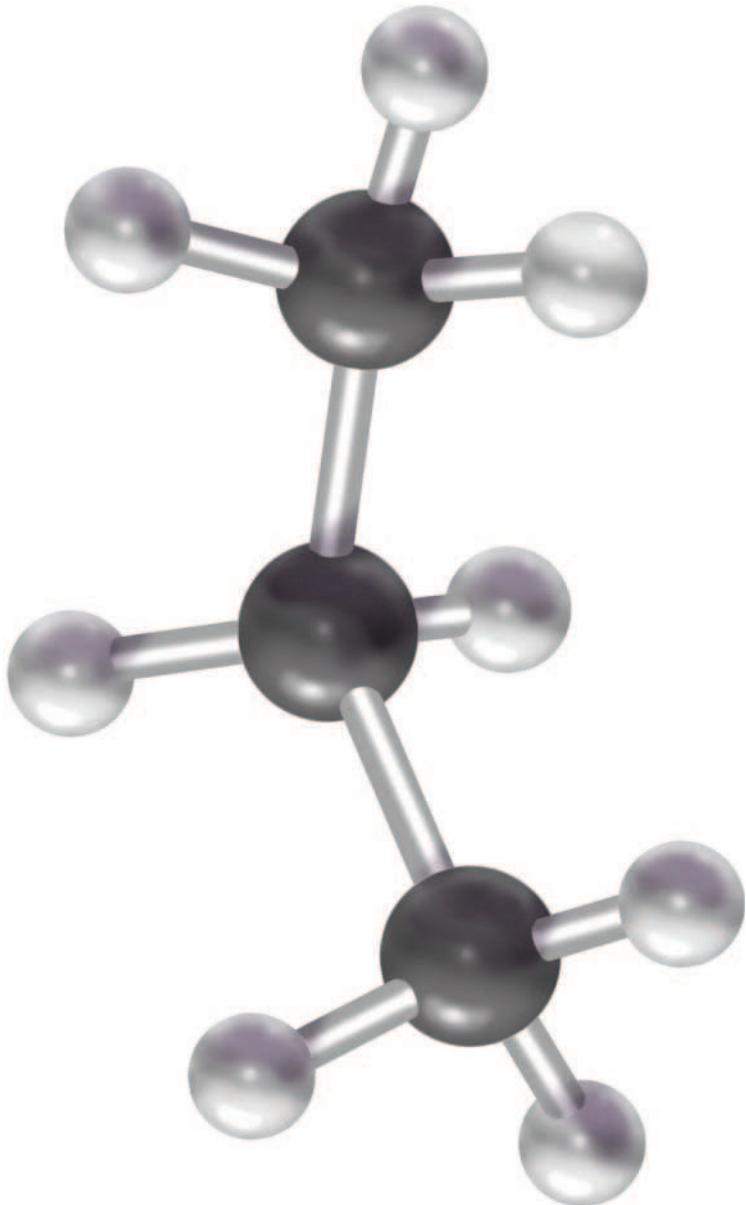
Propane (PRO-pane) is a colorless gas with an odor of natural gas. It occurs naturally in petroleum and natural gas. It belongs to the alkane family of organic compounds, compounds consisting of only carbon and hydrogen, all joined by single bonds. Propane is commonly sold as fuel, often available in a liquefied form known as liquid propane gas, or LPG.

HOW IT IS MADE

Propane is most widely available as a component of petroleum and natural gas, fossil fuels that formed many millions of years ago when marine organisms died, sank to the bottom of seas, and were eventually buried under massive layers of debris. The decay of those organisms without access to oxygen resulted in the formation of so-called fossil fuels: natural gas, petroleum, and coal. All fossil fuels are complex mixture of some free carbon and a very large variety of

PROPANE

Propane. White atoms are hydrogen; black atoms are carbon. PUBLISHERS RESOURCE GROUP



hydrocarbons. Natural gas, for example, consists primarily of methane, ethane, and propane.

Roughly half of all the propane produced in the United States comes from petroleum gases produced during the refining of crude oil and half from natural gas. The mixture

Interesting Facts

Propane's natural odor is so faint that it can often not be detected when leaks occur. To avoid the problem of unexpected fires and

explosions, manufacturers usually add a compound with a strong odor, such as ethanediol, to make leaks more noticeable.

of gases from either refined petroleum or natural gas is liquefied and then allowed to boil off, changing back to the original gases. Each gas boils off at its characteristic boiling point and can be captured and removed as it escapes from the liquid mixture. Often, propane is allowed to remain in its liquid state and is then made available as liquid propane gas (LPG). LPG is easier to store and transport than is gaseous propane.

COMMON USES AND POTENTIAL HAZARDS

About 100 billion liters (27 billion gallons) of propane were produced in the United States in 2004. The largest portion of that output (45 percent) was used for space heating, water heating, the operation of appliances, and other purposes in commercial and residential buildings. Propane has many advantages as a fuel:

- It produces less pollution than gasoline or coal.
- Propane water heaters are less expensive to purchase and operate than are electric water heaters, and they can heat more water than electric heaters.
- Propane fireplaces are less expensive and less polluting than wood-burning fireplaces, and they can be turned on and off with a switch.
- Many professional cooks prefer propane stoves to electric stoves because they heat instantly and are easier to control.
- Propane dryers take three-quarters of the time of electric dryers to dry the same amount of clothes.
- Propane appliances continue to operate during power outages, unlike electric appliances.

Words to Know

ALKANE One of a family of organic compounds consisting of only carbon and hydrogen, all joined by single bonds.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

The second most important application of propane is in the petrochemical industry where the compound is used for the production of a host of organic chemicals, primarily ethylene and propylene. Ethylene and propylene are two important compounds used in the production of plastics and other kinds of polymers. They ranked third and sixth on the list of the largest volume chemicals produced in the United States in 2004.

Farm use and industrial use follow next in importance, each accounting for about 7 percent of the propane used in the United States each year. On farms, propane is used to operate farm equipment and irrigation systems, to dry crops, and to help control weeds. In industries, the compound is used for space heating, for soldering, for cutting and treating metals, and as a fuel for fork lifts. Finally, a relatively small amount of propane is used as a fuel for cars, trucks, and other vehicles. Many people argue that more vehicles should be built to operate on propane because the gas produces fewer pollutants, reduces wear and tear on engines, and burns more efficiently than gasoline. So far, however, propane-powered vehicles have not become very popular in the United States. In the early 2000s, there were about 350,000 propane-powered vehicles in the United States and about 4 million in the world.

In moderate amounts, propane presents no health hazards to humans. In larger quantities, it may have a narcotic effect, producing drowsiness, disorientation, headaches, and confusion. The gas is both flammable and explosive and requires special care when being used.

FOR FURTHER INFORMATION

“Facts about Propane.” National Propane Gas Association.
http://www.npga.org/files/public/Facts_About_Propane.pdf
(accessed on November 3, 2005).

"Propane." Air Liquide.

http://www.airliquide.com/en/business/products/gases/gas_data/index.asp?GasID=53 (accessed on November 3, 2005).

"Propane Education and Research Council."

<http://www.propanecouncil.org/> (accessed on November 3, 2005).

"Propane: Exceptional Energy."

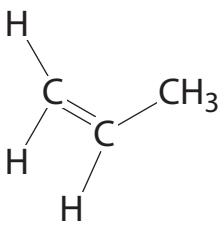
<http://www.usepropane.com> (accessed on November 3, 2005).

"Propane: What is Propane?" Vehicle Buyer's Guide for Consumers.

<http://www.eere.energy.gov/cleancities/vbg/consumers/lpg.shtml> (accessed on November 3, 2005).

Sherman, Josepha, and Steve Brick. *Fossil Fuel Power*. Mankato, Minn.: Capstone Press, 2003.

See Also Ethylene; Methane; Propylene

**OTHER NAMES:**

Propene; methylethylen; methylethene

FORMULA:

$\text{CH}_2=\text{CHCH}_3$

ELEMENTS:

Carbon, hydrogen

COMPOUND TYPE:

Alkene; unsaturated hydrocarbon (organic)

STATE:

Gas

MOLECULAR WEIGHT:

42.08 g/mol

MELTING POINT:

-185.24°C
 (-301.43°F)

BOILING POINT:

-47.69°C (-53.84°F)

SOLUBILITY:

Slightly soluble in water; very soluble in ethyl alcohol and ether

KEY FACTS

Propylene

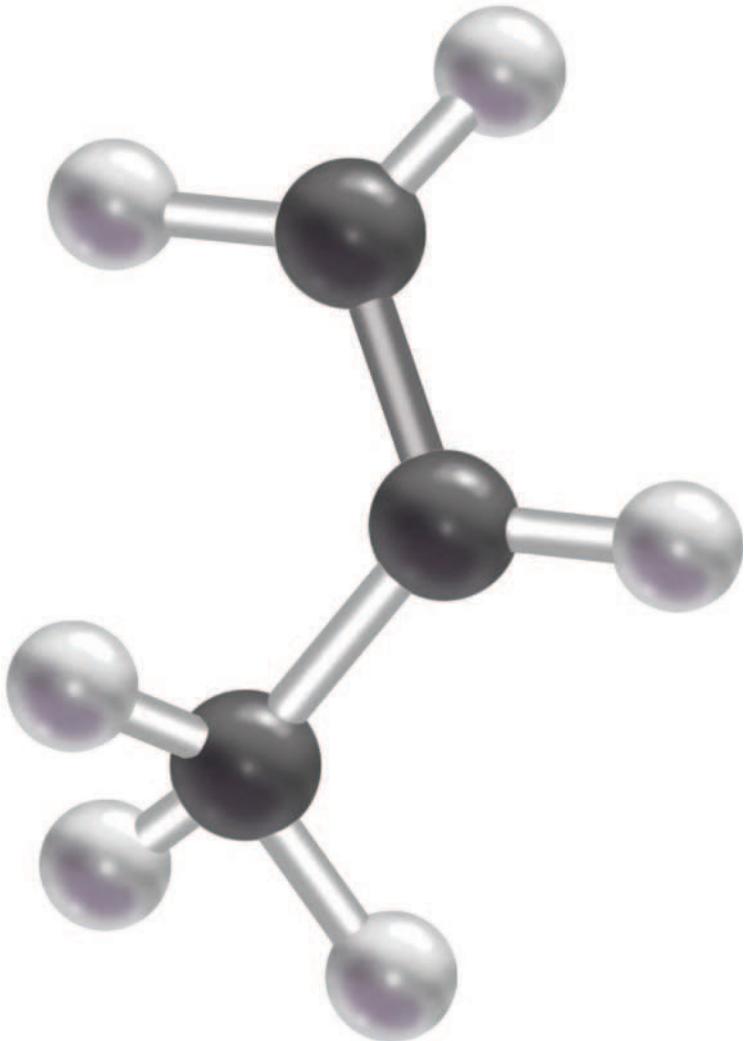
OVERVIEW

Propylene (PRO-puh-leen) is a colorless gas with a slightly sweet odor that burns with a yellow, sooty flame. It ranks sixth among all chemicals produced in the United States and second (after ethylene) among all organics. Its primary use is in the manufacture of polypropylene, one of the most popular polymers produced in the world.

HOW IT IS MADE

Propylene is prepared commercially by the thermal or catalytic cracking of hydrocarbons. The term cracking refers to a process by which large hydrocarbons—organic compounds that contain only carbon and hydrogen—are broken down into smaller hydrocarbons. Cracking can be accomplished either by exposing the hydrocarbons to high temperatures, a process called thermal cracking, or by using a catalyst, a process called catalytic cracking. A catalyst is a material that increases the rate of a chemical reaction without

Propylene. White atoms are hydrogen and black atoms are carbon. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP



undergoing any change in its own chemical structure. Thermal cracking is also known as steam cracking because steam is used to produce the high temperatures need to bring about the cracking reactions. Propylene can also be prepared by the catalytic dehydrogenation of propane (C_3H_8). Catalytic dehydrogenation is a process by which hydrogen atoms are removed from a substance, resulting in the formation of double bonds where single bonds previously existed.

Interesting Facts

- Plants emit small amounts of propylene gas naturally.
- Small amounts of propylene are also produced when organic matter burns and is found in products such as cigarette smoke, automobile exhaust, and burning leaves.

COMMON USES AND POTENTIAL HAZARDS

About 15.3 million metric tons (16.8 million short tons) of propylene were produced for commercial sale in the United States in 2004. About 39 percent of that amount was used for the production of polypropylene. Almost all of the remaining production was also used for the synthesis of chemical compounds, especially acrylonitrile (14 percent), propylene oxide (11 percent), cumene (10 percent), oxo alcohols (8 percent), isopropyl alcohol (7 percent), oligomers (5 percent) and acrylic acid (3 percent). Acrylonitrile, propylene oxide, oxo alcohols, and acrylic acids are all used primarily for the production of various types of polymers. Cumene is itself used as a raw material in the production of other organic compounds, especially acetone and phenol.

In addition to the propylene produced for commercial uses, large amounts of the gas are retained by the petroleum industry for conversion to products that are added to gasoline and other petroleum-based products. According to some estimates, between half and three-quarters of all propylene that comes from cracking reactions is retained for this purpose, leaving the remainder for commercial sale.

In addition to the hazard it poses as a flammable gas, propylene does have some health risks. At relatively high concentrations, it can act as an asphyxiant, a substance that can produce unconsciousness, and can cause death. It can also act as a mild anesthetic.

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

POLYMER A compound consisting of very large molecules made of one or two small repeated units called monomers.

FOR FURTHER INFORMATION

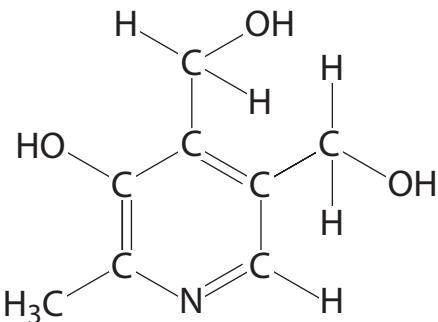
"Chronic Toxicity Summary: Propylene." California Office of Environmental Health Hazard Assessment.
http://www.oehha.org/air/chronic_rels/pdf/115071.pdf
(accessed on November 3, 2005).

Meikle, Jeffrey L. *American Plastic: A Cultural History*. Piscataway, NJ: Rutgers University Press, 1997.

"Propylene." Equistar Chemical Company.
<http://www.equistarchem.com/html/petrochemical/olefins/propylene.htm> (accessed on November 3, 2005).

"Propylene." LG Petrochemical.
http://www.lgpetro.com/eng/product_info/product/propylene.html (accessed on November 3, 2005).

See Also Ethylene; Polypropylene



OTHER NAMES:

See Overview.

FORMULA:

$\text{CH}_3\text{C}_5\text{HN(OH)(-CH}_2\text{OH)}_2$

ELEMENTS:

Carbon, hydrogen, nitrogen, oxygen

COMPOUND TYPE:

Organic

STATE:

Solid

MOLECULAR WEIGHT:

169.18 g/mol

MELTING POINT:

159°C-162°C (318°F-324°F)

BOILING POINT:

Not applicable; sublimes above melting point

SOLUBILITY:

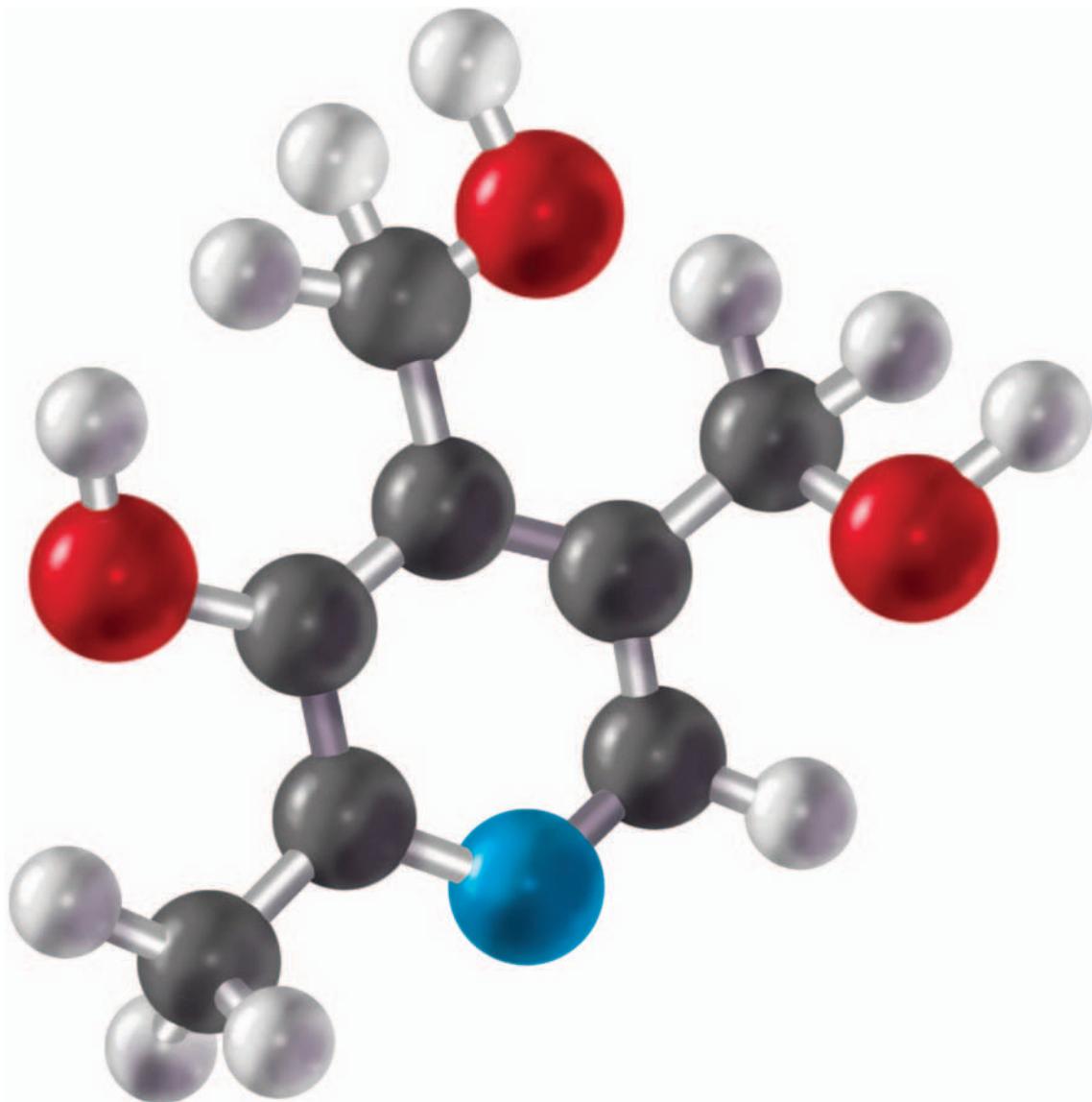
Very soluble in water; slightly soluble in ethyl alcohol and acetone

KEY FACTS

OVERVIEW

Pyridoxine (peer-ih-DOCK-seen) is also known as 3-Hydroxy-4,5-bis(hydroxymethyl)-2-methylpyridine; 3-hydroxy-4,5-dimethylol-2-methylpyridine; and vitamin B₆. It is a white, odorless, crystalline compound with a slightly bitter taste. The term pyridoxine is also used as a generic term for three compounds with biological activity classified under the term Vitamin B₆. The three compounds are pyridoxine, pyridoxal, and pyridoxamine. Pyridoxine is usually produced commercially as the hydrochloride, $\text{CH}_3\text{C}_5\text{HN(OH)(CH}_2\text{OH)}_2\cdot\text{HCl}$, which has somewhat different physical characteristics from pyridoxine itself.

Vitamin B₆ was discovered in 1938 by five groups of researchers working independently. The five groups were all looking for a cure for a disease in rats called acrodynia that resembles the human disease pellagra. In the early 1930s, the Hungarian-American biochemist Albert Szent-Györgyi (1893-1986) hypothesized the existence of a vitamin that would cure acryodynia and even gave the vitamin a



Pyridoxin. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atom is nitrogen. Ring of alternating white and dark sticks is a benzene ring. PUBLISHERS RESOURCE GROUP

name, vitamin B₆. So the simultaneous discovery of such a compound in 1938 was not much of a surprise.

The first component of vitamin B₆, pyridoxine, was first synthesized, also in 1938, by the Austrian-German chemist Richard Kuhn (1900-1967). Its chemical structure was determined a year later by American chemists Karl August Folkers (1906-1997) and S. A. Harris (dates not available) at the

Interesting Facts

Because pyridoxine is water soluble, it dissolves when foods are cooked or processed.

Merck chemical corporation. Pyridoxine is the most stable form of the vitamin, so it is the form used in vitamin supplements and as a food additive.

HOW IT IS MADE

Pyridoxine is produced naturally by most plants and animals in sufficient amounts to prevent vitamin B₆ deficiency diseases. It is also produced synthetically by a complex series of reactions that begins with isoquinoline (C₉H₇N).

COMMON USES AND POTENTIAL HAZARDS

Vitamin B₆ has been shown to be essential in many biochemical reactions that occur in plants and animals. Although it may occur in any one of the three forms listed above, the compound usually acts as the phosphate ester, pyridoxine phosphate. Pyridoxine phosphate functions as a coenzyme in the transformation of amino acids, the building blocks from which proteins are made. A coenzyme is a chemical compound that works with an enzyme to catalyze some essential chemical reaction in the body. Pyridoxine phosphate appears to be necessary for the synthesis of proteins from amino acids as well as the metabolism of amino acids to produce energy needed for normal body functioning.

Vitamin B₆ deficiency diseases are very rare. In 1954, a batch of commercially prepared baby food was overheated during its preparation. Overheating apparently destroyed the vitamin B₆ present in the food. Babies who were fed with the food had convulsions became unusually irritable, and developed unusual behaviors. As soon as the babies were given vitamin B₆ supplements, these symptoms disappeared. Such instances among humans are so rare that they become the

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

METABOLISM The process including all of the chemical reactions that occur in cells by which fats, carbohydrates, and other

compounds are broken down to produce energy and the compounds needed to build new cells and tissues.

SYNTHESIS Chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

subject of articles in medical journals. Other reported instances of vitamin B₆ deficiency disease have involved pregnant women who did not receive enough of the vitamin in their daily diets and people living in Cuba during the early 1990s who had restricted diets. Symptoms of vitamin B₆ deficiency include general weakness, anemia, cracked lips, inflamed tongue and mouth, irritability, depression, and skin disorders.

Meats have the highest concentration of vitamin B₆, so vegetarians may be at risk for deficiency disorders. Other foods that contain high concentrations of the vitamin include bananas, mangoes, avocados, and potatoes. Increased doses of vitamin B₆ are sometimes used to treat morning sickness and insomnia, and some authorities recommend the vitamin to decrease the risk of heart disease. The maximum recommended dose of vitamin B₆ is 50 milligrams a day.

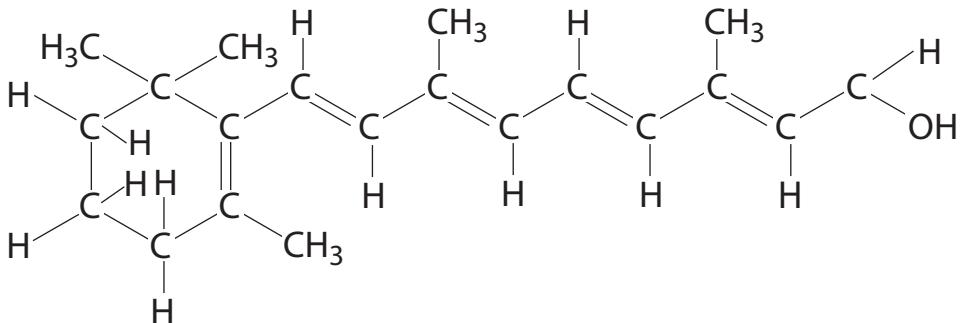
FOR FURTHER INFORMATION

Brody, Tom. *Nutritional Biochemistry*. San Diego: Academic Press, 1998.

"Dietary Supplement Fact Sheet: Vitamin B₆." NIH Office of Dietary Supplements.

<http://ods.od.nih.gov/factsheets/vitaminb6.asp> (accessed on November 3, 2005).

Turner, Judith. "Pyridoxine." In *Gale Encyclopedia of Alternative Medicine*. Detroit: Gale Group, 2004.



OTHER NAMES:

Vitamin A

FORMULA:

$C_{20}H_{30}O$

ELEMENTS:

Carbon, hydrogen,
oxygen

COMPOUND TYPE:

Alcohol (organic)

STATE:

Solid

MOLECULAR WEIGHT:

286.45 g/mol

MELTING POINT:

63.5°C (146°F)

BOILING POINT:

Not applicable;
decomposes

SOLUBILITY:

Practically insoluble
in water; soluble in
ethyl alcohol and
methyl alcohol

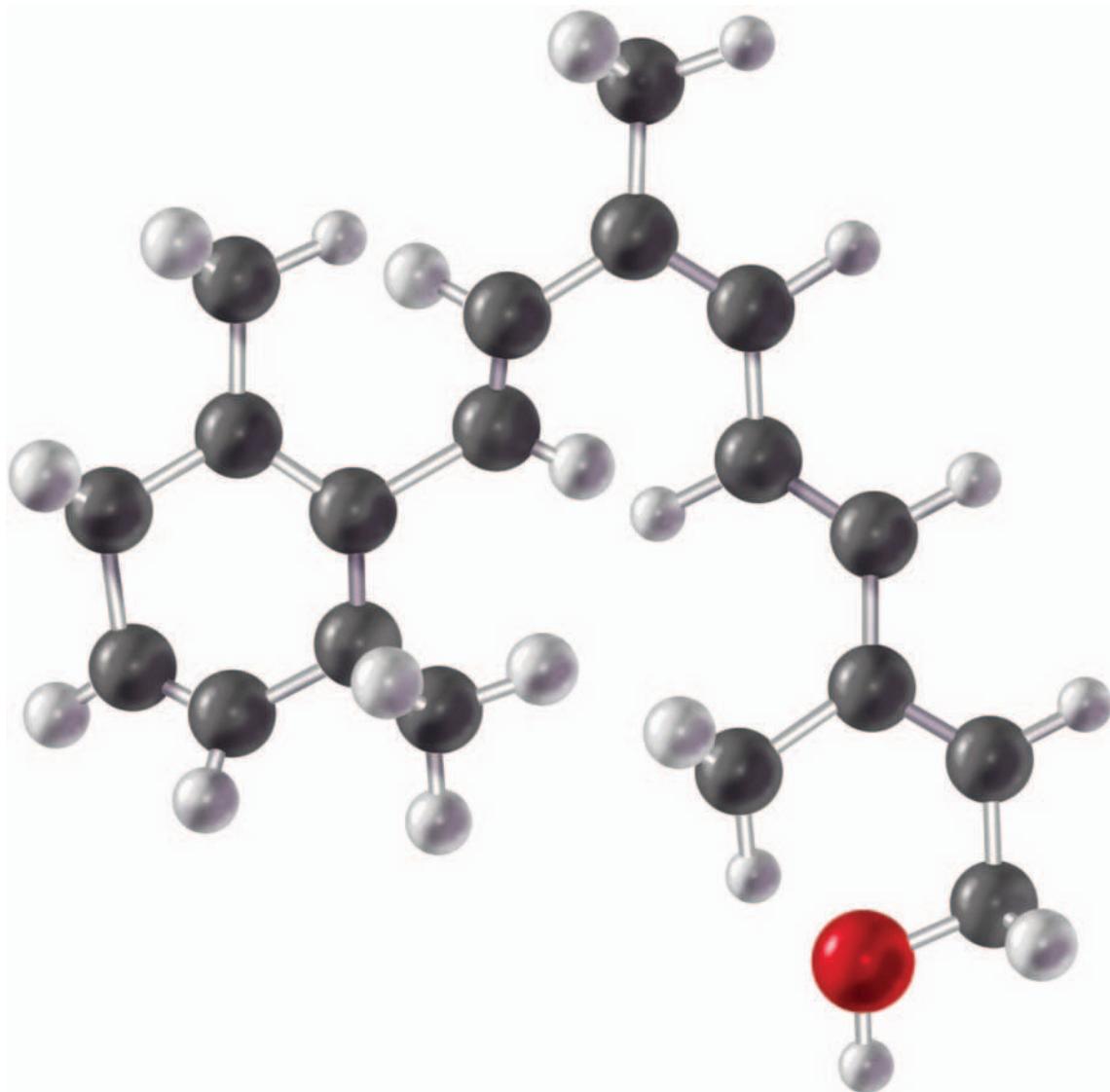
K E Y F A C T S

OVERVIEW

Retinol (RET-uh-nol) is the scientific name for vitamin A, a vitamin found only in animals. It occurs as a yellowish to orange powder with a slight brownish cast and is a relatively stable compound. Retinol is converted in the body from an alcohol to the corresponding aldehyde, retinal ($C_{20}H_{28}O$), one of the primary chemical compounds involved in the process by which light is converted to nerve impulses in the retina of the eye. Vitamin A is also required for a number of other biochemical reactions in the body, including growth and development of tissue and maintenance of the immune system.

HOW IT IS MADE

Vitamin A is synthesized in animal bodies through a variety of pathways. One important source of vitamin A is a group of related compounds called the carotenes, substances responsible for the yellowish or orangish appearance of



Retinol. Red atom is oxygen; white atoms are hydrogen; and black atoms are carbon. Gray sticks indicate double bonds.

PUBLISHERS RESOURCE GROUP

fruits and vegetables such as carrots, sweet potatoes, squash, cantaloupe, apricots, pumpkin, and mangos. Some leafy green vegetables, such as collard greens, spinach, and kale, are also good sources of the carotenes. The most important of the carotenes is β -carotene (beta-carotene), $C_{40}H_{56}$. The oxidation of carotenes in animal bodies converts them to retinol.

The chemical structure of retinol was determined in 1931 by Swiss chemist Paul Karrer (1889–1971), and the compound

Interesting Facts

- Animals that live in very cold climates have high concentrations of retinol in their livers. A polar bear liver contains enough retinol to kill a human

who eats it. Vitamin A deficiency is the leading cause of childhood blindness in the world today.

was first prepared synthetically shortly thereafter by Austrian-German chemist Richard Kuhn (1900–1967). The first successful process for producing retinol commercially was developed in the mid-1940s by German chemist Otto Isler (1920–1992), then employed at the pharmaceutical company Roche, located in Sissein, Germany. Isler's process involved a complex series of reactions that begins with the combination of a fourteen carbon hydrocarbon and a six carbon hydrocarbon to create the fundamental backbone from which the retinol molecule is constructed. Regular production of vitamin A began in 1948 with a projected output of 10 kilograms per month, which before long was raised to 50 kilograms per month. The Roche plant at Sissein continues to produce retinol today.

COMMON USES AND POTENTIAL HAZARDS

Vitamin A is probably best known for its role in maintaining normal vision. Deficiencies of the compound are likely to manifest themselves earliest in a variety of eye problems, most commonly night blindness. Night blindness is a condition in which one loses the ability to distinguish objects in reduced light. If left untreated, vitamin A deficiencies may lead to decreased ability to see in normal light and, eventually, to complete blindness.

But vitamin A has been shown to have a number of other functions in the body. It is essential for the maintenance of growth, bone formation, reproduction, proper immune system function, and healing of wounds. A number of additional claims have been made for the compound, although evidence is not as strong as it is for the above functions. For example,

it may be effective in preventing or treating a variety of conditions such as measles, intestinal parasites, osteoporosis, inflammatory bowel disease, bone marrow disorders, certain types of cancer, tuberculosis, peritonitis, osteoarthritis, food poisoning, Alzheimer's disease, miscarriage, and HIV/AIDS. In each of these cases, evidence is not yet strong enough to show a clear-cut connection between retinol and disease, but research is being conducted to determine how strong the association may be.

Retinol is available commercially in a variety of formulations, including tablets, capsules, and creams. Such products usually contain a modified form of retinol that is more easily absorbed by the body. For example, a product known as tretinoin is a synthetic form of retinol known as all-trans retinoic acid. The term *all trans* means that all of the double bonds in retinoic acid are located on the same side of the molecule. Products containing tretinoin are used to treat acne, pimples, wrinkles, blackheads, freckles, sun-spots, and even pre-cancerous lesions. They work by increasing the rate with which the skin sheds old cells and replaces them with new cells.

Vitamin A supplements in pill or capsule form are available in two formulations, those that contain retinol and those that contain beta carotene. It is not possible to take too much of the latter type of vitamin A. The body will not convert excess amounts of carotene into retinol but will, instead, excrete the excess in the urine or stool. An excess of retinol-based vitamin A, by contrast, may result in certain medical problems. Since the vitamin is fat soluble, it may be stored in body fat and reach relatively high concentrations if too much is ingested. An excess of retinol in the body may be associated with liver damage, osteoporosis, rash, fatigue, bone and joint pain, nausea, insomnia, and personality changes.

FOR FURTHER INFORMATION

Powers, Jennifer I. "Acne Medication." In *Chemistry: Foundations and Applications*. Volume 1. Edited by J. J. Lagowski. New York: Macmillan Reference USA, 2004, 12 15.

"Vitamin A." Bristol University School of Chemistry.
<http://www.chm.bris.ac.uk/webprojects2002/schnepf/vitaminA.html> (accessed on November 3, 2005).

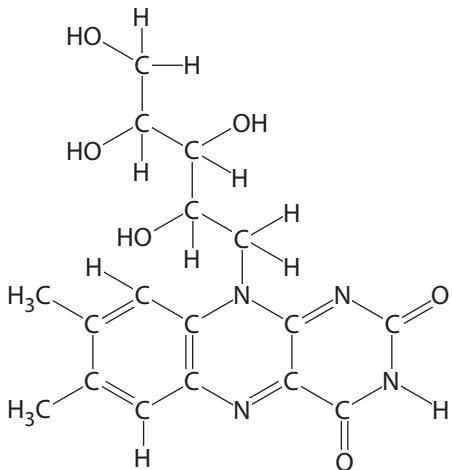
"Vitamin A (Retinol)." Hypertexts for Biomedical Sciences.

http://arbl.cvmbs.colostate.edu/hbooks/pathphys/misc_topics/vitamina.html (accessed on November 3, 2005).

"Vitamin A (Retinol)." University of Maryland Medical Center.

http://www.umm.edu/altmed/ConsSupplements/VitaminA_Retinolcs.html (accessed on November 3, 2005).

See Also Beta-Carotene

**OTHER NAMES:**Vitamin B₂**FORMULA:**C₁₇H₂₀N₄O₆**ELEMENTS:**

Carbon, hydrogen, nitrogen, oxygen

COMPOUND TYPE:

Organic

STATE:

Solid

MOLECULAR WEIGHT:

376.36 g/mol

MELTING POINT:

280°C (536°F)

BOILING POINT:

Not applicable; decomposes

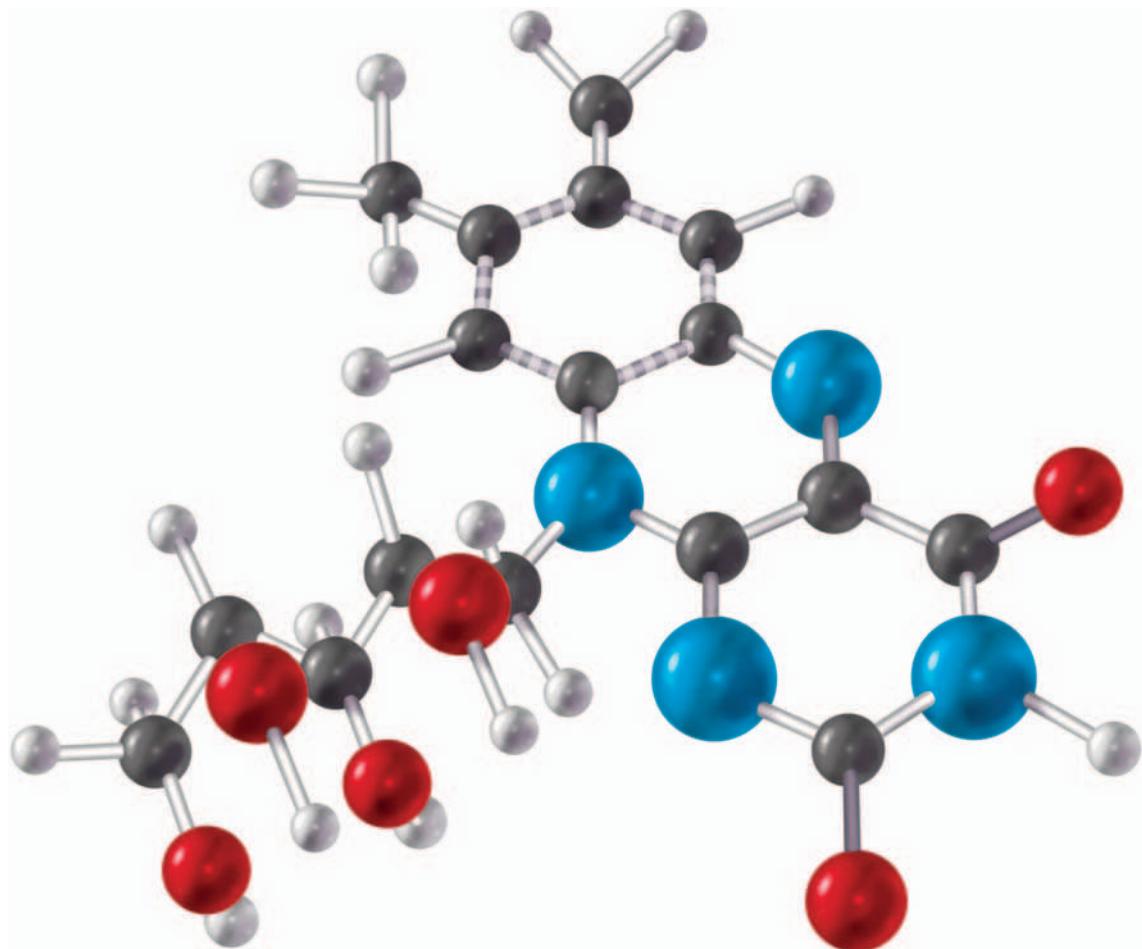
SOLUBILITY:

Very slightly soluble in water and ethyl alcohol; very soluble in alkaline solvents, but resulting in decomposition

KEY FACTS**OVERVIEW**

Riboflavin (REY-bo-FLAY-vin), commonly known as vitamin B₂, is an orange-yellow crystalline solvent with a bitter taste. It is relatively stable when exposed to heat, but tends to decompose in the presence of light for extended periods of time. Riboflavin is used in the body for a variety of functions, including the metabolism of carbohydrates for the production of energy and the production of red blood cells.

Riboflavin was first observed in 1879 by the English chemist Alexander Wynter Blyth (1844-1921) who noticed a compound in cow's milk that glowed with a yellow fluorescence when exposed to light. Blyth called the compound *lachtochrome* (*lachto-* = "milk" and *-chrome* = color), but was unable to determine its chemical composition or its chemical properties. In fact, it was not until the 1930s that the chemical nature of the compound was determined. The Swiss chemist Paul Karrer (1889-1971) and the Austrian-German chemist Richard Kuhn (1900-1967) independently determined the chemical structure of riboflavin and first



Riboflavin. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atoms are nitrogen. Gray sticks indicate double bonds. Striped sticks show a benzene ring,

PUBLISHERS RESOURCE GROUP

synthesized the compound. The name *riboflavin* is derived from the fact that the vitamin was first found in association with the sugar ribose.

HOW IT IS MADE

Plants and microorganisms synthesize riboflavin naturally. Some foods rich in riboflavin are brewer's yeast, dark green vegetables, mushrooms, legumes, nuts, milk and other dairy products, sweet potatoes, and pumpkins. Bacteria that live in the human digestive tract are also able to synthesize some riboflavin, but not enough to meet the body's requirement for the vitamin.

Interesting Facts

- The human body cannot store riboflavin, so it is excreted in the urine. For this reason, it is not dangerous to consume large doses of riboflavin, although the consumption of large amounts of the vitamin serves no biological purpose.
- Since riboflavin does not dissolve in water, it must be converted to a water soluble form, such as riboflavin-5-phosphate, for use as a food additive.
- The recommended daily dose of riboflavin for adults is 1.6 milligrams for men, 1.2 milligrams for women, 0.6 milligrams for children age four to eight, 0.9 milligrams for children age nine to thirteen, and 1.3 milligrams for teenagers.

Riboflavin is produced synthetically using either the genetically-modified bacterium *Bacillus subtilis* or a fungus called *Ashbya gossypii*. The bacteria or fungus are cultured in a large vat that has been seeded with small amounts of riboflavin. Over time, the organisms generate large quantities of riboflavin until some desired amount of the compound has been produced. The vat is then heated to a temperature sufficient to kill the bacteria or fungi, leaving crystalline riboflavin behind. The riboflavin is then separated and purified.

COMMON USES AND POTENTIAL HAZARDS

The human body needs riboflavin to use oxygen efficiently in the metabolism of amino acids, fatty acids, and carbohydrates. The vitamin is involved in the synthesis of niacin (another B vitamin), it activates vitamin B₆, and it helps the adrenal gland to produce hormones. It helps the body make antibodies to fight disease and infection, regulates the thyroid gland, and is important in maintaining healthy hair, nails, and skin. Riboflavin is especially important during periods of rapid growth because it is involved in the formation and growth of cells, especially red blood cells.

Words to Know

CARBOHYDRATE A group of sugars and starches produced by plants and used as food.

JAUNDICE A disease caused by malfunction of the liver in which the skin and eyes become yellow.

LEGUME A member of the pea family of plants, which includes peas and beans.

METABOLISM The process that includes all of the chemical reactions that occur in cells by which fats, carbohydrates, and other compounds are broken down to

produce energy and the compounds needed to build new cells and tissues.

MUCOUS MEMBRANES Tissues that line the moist inner lining of the digestive, respiratory, urinary and reproductive systems.

SOLVENT A substance that is able to dissolve one or more other substances.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

In combination with vitamin A, riboflavin helps maintain the mucous membranes that line the digestive tract. Pregnant women need riboflavin to help the fetus grow and develop. The vitamin is also essential for eye health. Some medical professionals recommend riboflavin for the treatment of eye disorders, such as cataracts, sensitivity to bright light, and bloodshot, burning, or itching eyes. Doctors have begun experimenting with the use of riboflavin to treat migraine headaches.

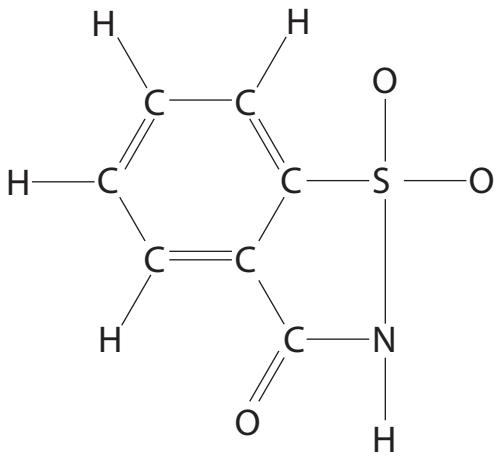
In spite of its many important functions in the body, riboflavin deficiencies do not lead to serious medical problems or death. They may result in a delayed healing of wounds or relatively minor medical conditions such as cheilosis (a reddening and soreness in the corners of the lips), angular stomatitis (cracking of the skin at the corner of the lips), dermatitis (an oily skin disorder), glossitis (a swollen, reddened tongue), and cracking around the nose. In all such cases, however, a riboflavin deficiency by itself is almost never the cause of the problem; instead, deficiencies of other vitamins in the B group are also involved.

Riboflavin deficiencies are very rarely seen in the United States because the vast majority of people consume a diet that contains adequate amounts of the vitamin. Individuals

most likely to suffer from riboflavin deficiency problems are those with anorexia (a condition in which people refuse to eat adequate amounts of food), older people with poor diets, alcoholics (because alcohol impairs a person's ability to absorb and use the vitamin), and newborn babies being treated for jaundice by exposure to ultraviolet light (because light destroys riboflavin).

FOR FURTHER INFORMATION

- "Food Safety: From the Farm to the Fork." European Commission Web Site.
http://europa.eu.int/comm/food/fs/sc/scf/out18_en.html
(accessed on November 3, 2005).
- "Riboflavin." Linus Pauling Institute, Micronutrient Information.
<http://lpi.oregonstate.edu/infocenter/vitamins/riboflavin/>
(accessed on November 3, 2005).
- "Riboflavin." MedlinePlus.
<http://www.nlm.nih.gov/medlineplus/ency/article/002411.htm> (accessed on November 3, 2005).
- "Vitamin B₂ (Riboflavin)." Herb & Supplement Encyclopedia.
<http://www.florahealth.com/flora/home/canada/healthinformation/encyclopedia/VitaminB2.asp> (accessed on November 3, 2005).



OTHER NAMES:	Benzoylsulfonic imide; benzoic sulfimide
FORMULA:	$C_7H_5NO_3S$
ELEMENTS:	Carbon, hydrogen, nitrogen, oxygen, sulfur
COMPOUND TYPE:	Organic
STATE:	Solid
MOLECULAR WEIGHT:	183.18 g/mol
MELTING POINT:	228°C (442°F)
BOILING POINT:	Not applicable; decomposes
SOLUBILITY:	Slightly soluble in water; soluble in acetone and ethyl alcohol

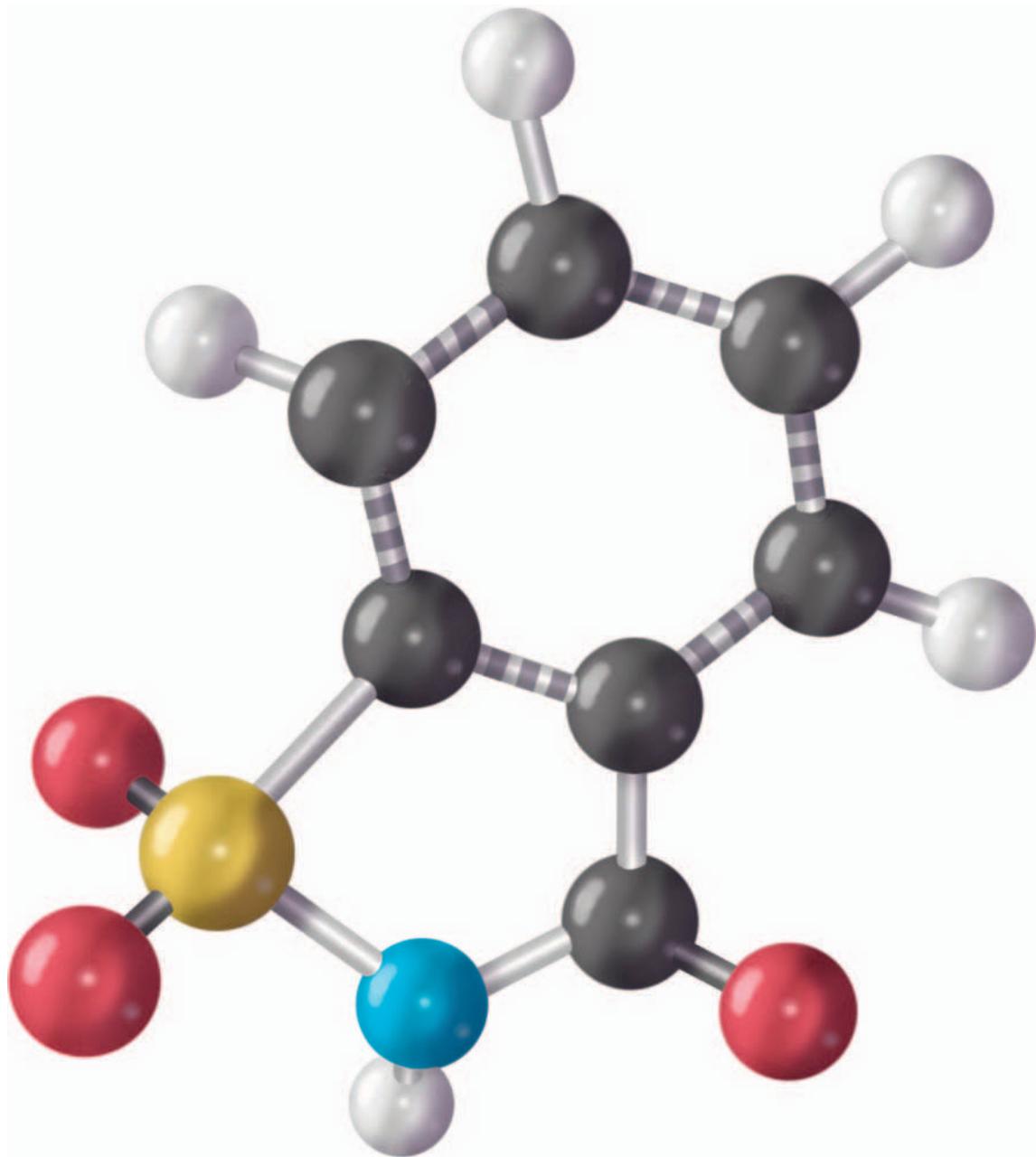
KEY FACTS

Saccharin

OVERVIEW

Saccharin (SAK-uh-rin) is a synthetic compound whose water solutions are at least 500 times as sweet as table sugar. It passes through the human digestive system without being absorbed, so it has an effective caloric value of zero. It is used as a sugar substitute by diabetics or by anyone wishing to reduce their caloric intake.

Saccharin was the first artificial sweetener discovered. It was synthesized accidentally in 1879 when Johns Hopkins researchers Constantine Fahlberg (1850-1910) and Ira Remsen (1846-1927) were working on the development of new food preservatives. The story is told that Fahlberg accidentally spilled one of the substances being studied on his hand. Some time later, he noticed the sweet taste of the substance and began to consider marketing the product as an artificial sweetener. Fahlberg and Remsen jointly published a paper describing their work, but Fahlberg, without Remsen's knowledge, went on to request a patent for the discovery. He eventually became very wealthy from proceeds of the discovery, none of



Benzoic sulfinide. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; blue atom is nitrogen; and yellow atom is sulfur. Gray sticks indicate double bonds. Striped sticks show a benzene ring.

PUBLISHERS RESOURCE GROUP

which he shared with Remsen. Remsen was later quoted as saying that “Fahlberg is a scoundrel. It nauseates me to hear my name mentioned in the same breath with him.”

In spite of its sweet taste, saccharin was used at first as an antiseptic, a substance that stops the growth of or kills bacteria. Before long, however, its primary use became that

of an artificial sweetener. By 1902, it had become so popular in Germany that the German sugar industry lobbied for laws limiting production of saccharin. Similar actions occurred in the United States in 1907 and, by 1911, the federal government restricted use of the compound to overweight invalids.

A shortage of sugar during World War I (1914–1918) led to the reintroduction of saccharin as a sweetening agent in foods. Another sugar shortage during World War II (1939–1945) saw a new boom in saccharin production. This time, the compound's popularity continued after the war ended.

Questions about saccharin's safety have been raised a number of times in the past. In 1969, for example, the sweetener cyclamate was found to be carcinogenic, and its use was banned in the United States. Doubts over saccharin's safety resurfaced, partly since it was often mixed with cyclamate in artificial sweeteners. In 1972, studies with rats suggested that saccharin too might be carcinogenic, and the U.S. Food and Drug Administration imposed restrictions on the sweetener's use. Later studies attempting to reproduce the 1972 research were largely unsuccessful, and the status of saccharin as a carcinogen remain unsettled.

In 1977, the Canadian government decided that sufficient evidence existed to ban the use of saccharin except for use with diabetics and others with special medical problems. The U.S. government considered taking similar action, but, after more than a decade of reviewing the evidence, decided to allow the use of saccharin among the general public. Nonetheless, the status of saccharin as a potential health hazard remains the subject of an active debate in the United States and other parts of the world.

HOW IT IS MADE

A number of methods are available for the synthesis of saccharin. For many years, the most popular process was one developed by the Maumee Chemical Company of Toledo, Ohio, in 1950. This method begins with anthranilic acid (*o*-aminobenzoic acid; $C_6H_4(NH_2)COOH$), which is treated successively with nitrous acid (HNO_2), sulfur dioxide (SO_2), chlorine (Cl_2), and ammonia (NH_3) to obtain saccharin. Another process discovered in 1968 starts with *o*-toluene, which is then treated with sulfur dioxide and ammonia to obtain saccharin.

Interesting Facts

- The chemical company Monsanto was founded in 1901 for the sole purpose of making saccharin. The company sold all of its

product to a single company, the Coca-Cola company, founded in the late 1880s.

Saccharin is not very soluble, so it is commonly made into its sodium or calcium salt (sodium saccharin or calcium saccharin), both of which readily dissolve in water, for use in drinks and cooking. Saccharin is often blended with other sweeteners to reduce its metallic aftertaste.

COMMON USES AND POTENTIAL HAZARDS

Saccharin is used almost exclusively as an artificial sweetener in food and drinks to replace sugar. Its lack of calories makes it suitable for diet products and for medical preparations designed for people who must reduce their caloric intake. It also finds some small application as a food preservative, as an antiseptic agent, and as a brightening agent in electroplating procedures.

Raw saccharin can be an irritant to the skin, eyes, and respiratory system. If ignited, it burns with the release of irritating fumes. Only individuals who come into contact

Words to Know

CARCINOGEN A substance that causes cancer in humans or other animals.

ELECTROPLATING A process by which a thin layer of one metal is deposited on top of a second metal by passing an electric current through a solution of the first metal.

SALT A compound in which all of the hydrogen ions are replaced by metal ions.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

with large quantities of saccharin are likely to be concerned about such safety problems, however.

FOR FURTHER INFORMATION

Henkel, John. "Sugar Substitutes: Americans Opt for Sweetness and Lite." *FDA Consumer*. December 1999. Available online at http://www.fda.gov/fdac/features/1999/699_sugar.html (accessed on November 3, 2005).

Nabors, Lyn O'Brien, ed. *Alternative Sweeteners (Food Science and Technology)*. Third revision. London: Marcel Dekker, 2001.

Ruprecht, Wilhelm. "Consumption of Sweeteners: An Evolutionary Analysis of Historical Development." <http://www.druid.dk/conferences/nw/paper1/ruprecht.pdf> (accessed on November 3, 2005).

"Saccharin Sodium." J. T. Baker.
<http://www.jtbaker.com/msds/englishhtml/soo73.htm>
(accessed on November 3, 2005).

**OTHER NAMES:**

Silica, quartz, sand, amorphous silica, silica gel, and others

FORMULA:

SiO_2

ELEMENTS:

Silicon, oxygen

COMPOUND TYPE:

Nonmetallic oxide
(inorganic)

STATE:

Solid

MOLECULAR WEIGHT:

60.08 g/mol

MELTING POINT:

Varies depending on crystalline state; typically above 1700°C (3100°F)

BOILING POINT:

2950°C (5300°F)

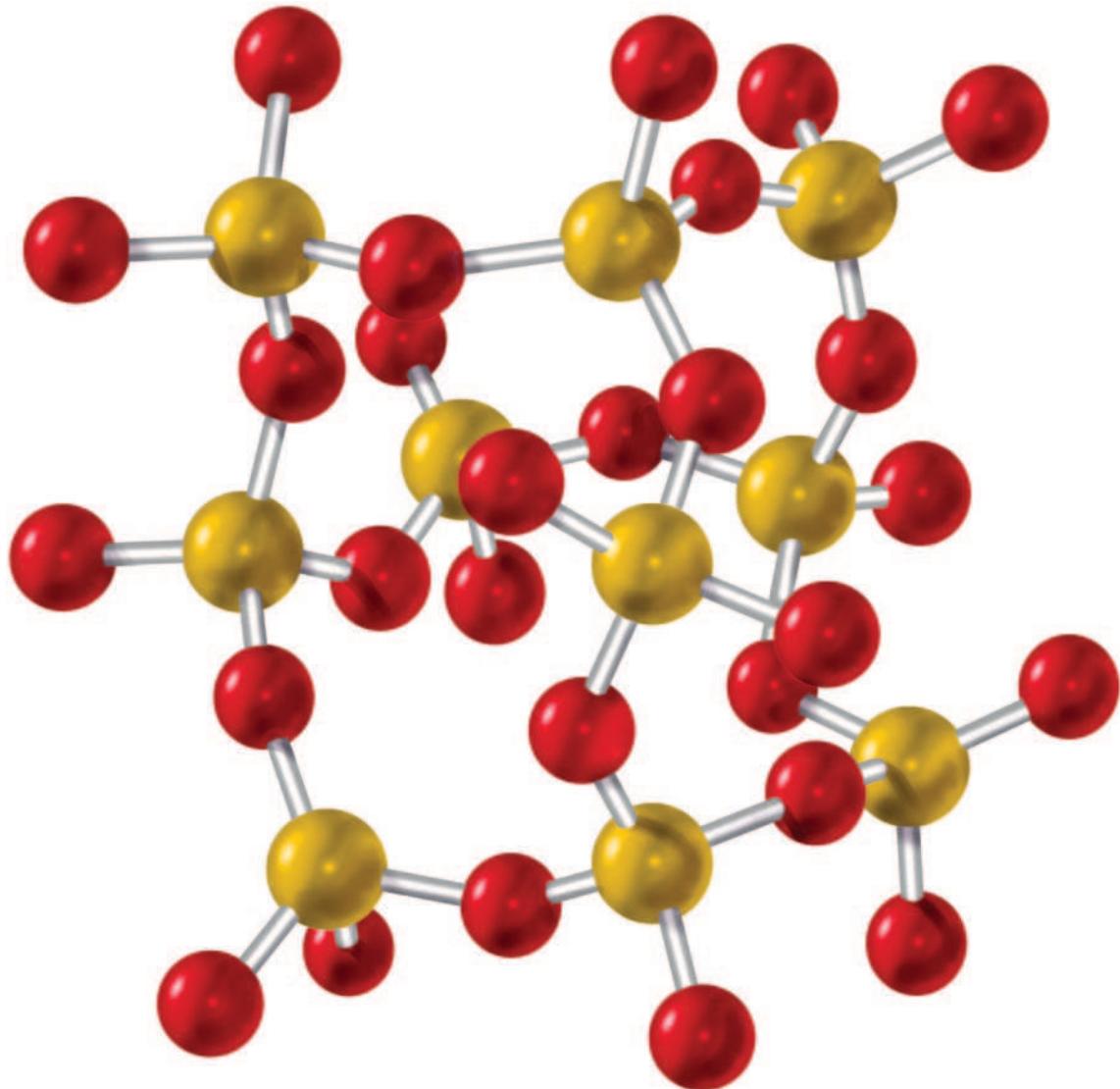
SOLUBILITY:

Solubility depends on crystalline state; generally insoluble in water; soluble in many acids and alkalis

KEY FACTS**OVERVIEW**

Silicon dioxide (SILL-uh-kon dye-OK-side) is one of the most abundant chemical compounds on Earth. It makes up about 60 percent of the weight of the Earth's crust either as an independent compound (SiO_2) or in combination with metallic oxides that form silicates. Silicates are inorganic compounds whose negative part is the SiO_3 ion (grouping of atoms). An example is magnesium silicate, MgSiO_3 .

Silicon dioxide occurs as colorless, odorless, tasteless white or colorless crystals or powder. Its many different forms can be classified as crystalline, amorphous, or vitreous. In crystalline forms of silicon dioxide, all of the atoms that make up the substances are arranged in orderly patterns that have the shape of cubes, rhombohedrons, or other geometric figures. In amorphous silicon dioxide, silicon and oxygen atoms are arranged randomly, without any clear-cut pattern. Vitreous silicon dioxide is a glassy form of the compound that may be transparent, translucent, or opaque. The various forms of silicon dioxide can be converted from one form to another by heating and changes in pressure.



Silicon dioxide. Red atoms are oxygen and yellow atoms are silicon. PUBLISHERS RESOURCE GROUP

An especially interesting form of silicon dioxide is silica gel, a powdery form of amorphous silicon dioxide that is highly adsorbent. An adsorbent material (in contrast to an absorbent material) is one that is capable of removing a material, such as water, ammonia, alcohol, or other gases, out of the air. The second material bonds weakly to the outer surface of silica gel particles. Silica gel is able to adsorb anywhere from 30 to 50 percent of its own weight in water.

Interesting Facts

- Stardust, a U.S. National Aeronautics and Space Administration (NASA) spacecraft, used silica gel to collect particles of debris from the tail of comet Wild-2.
- Although silica gel has been known since the mid-seventeenth century, practical applications for the material were not discov-

ered until 1919 when American Walter A. Patrick (1888–1969) patented a number of processes for the manufacture of the compound. It still did not become widely popular until World War II (1939–1945), when the American military found a number of important uses for the compound.

from the surrounding atmosphere before it becomes saturated. The silica gel is not chemically altered by the process of adsorption and still feels dry even when saturated. The adsorbed water can be driven off simply by heating the silica gel, allowing the material to regain its adsorbent properties.

HOW IT IS MADE

Although methods are available for synthesizing silicon dioxide, there is no practical reason for doing so. The abundant quantities of silicon dioxide found in the earth's crust are sufficient to satisfy all industrial needs. Among the minerals and earths that contain silicon dioxide in an uncombined form are quartz, flint, diatomite, stishovite, agate, amethyst, chalcedony, cristobalite, and tridymite.

Naturally occurring silicon dioxide can be treated by a variety of physical processes to change its form. For example, heating crystalline silicon dioxide above its melting point and then cooling it again converts the compound into its vitreous form, sometimes called natural glass. Silica gel is made by treating sodium silicate (Na_2SiO_3) with sulfuric acid (H_2SO_4): $\text{Na}_2\text{SiO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{SiO}_2 + \text{H}_2\text{O} + \text{Na}_2\text{SO}_4$.

COMMON USES AND POTENTIAL HAZARDS

The primary use of silicon dioxide is in the building industry. It is used to make ceramics, enamels, concrete, and specialized silica bricks used as refractory materials. It is also one of the raw materials from which all kinds of glass are made. Vitreous silicon dioxide is an important constituent of specialized types of glass, such as that used in making laboratory equipment, mirrors, windows, prisms, cells, and other kinds of optical devices. Silicon dioxide is also used as an anti-caking or thickening agent in a variety of foods and pharmaceutical products. Some other applications of silicon dioxide include:

- In the manufacture of polishing and grinding materials;
- As molds for casting;
- In the production of elemental silicon;
- As a filler in many different kinds of products, including paper, insecticides, rubber products, pharmaceuticals, and cosmetics;
- As an additive in paints to produce a low-gloss finish;
- In the reinforcement of certain types of plastics.

The primary application of silica gel is as a drying agent. Packets of silica gel are found in many consumer products, such as electronic equipment, hardware tools, clothing, CD and DVD discs, and foodstuffs. Because of its ability to adsorb moisture from the surrounding air, silica gel prevents rust and other forms of oxidation. Silica gel also has similar applications in industry. For example, it is used to dry compressed air, air conditioning systems, and natural gas. The compound is also used to bleach petroleum oils and as an anti-caking agent for cosmetics and pharmaceuticals.

Normal exposure to silicon dioxide is not considered hazardous to the health of humans or other animals. Inhalation of silicon dioxide dust or fumes, however, may create problems in the respiratory system. In low doses, inhaled silica can cause coughing, wheezing, and difficulty breathing. In higher doses, particles of silicon dioxide may lodge in the lungs and block the openings through which oxygen is absorbed. Over long periods of exposure, a condition known as silicosis may develop. Silicosis is a condition similar to tuberculosis, lung cancer, or emphysema in which a person's ability to breathe normally is reduced, causing severe and

Words to Know

ALKALI A strong base.

often used to line the interior of industrial furnaces.

REFRACTORY A material with a high melting point, resistant to melting,

life-threatening long-term health problems. Individuals at greatest risk for silicosis and other silicon dioxide-related problems are those who cut, chip, drill, or grind objects that contain silica. During these processes, silica is reduced to a fine powder that is easily inhaled. Wearing a mask during these operations is generally sufficient to protect a worker from inhaling these fumes and particles.

FOR FURTHER INFORMATION

Brown, David. ‘Walter Patrick: Preservationist of the First Order.’ *Mount Washington Newsletter*. Spring 2003. Also available online at http://www.mwia.org/Newsletters/MWINewsletter_Spring%202003.pdf (accessed on December 10, 2005).

Brownlee, Don. “An Exciting Encounter with a Cold Dark Mysterious Body from the Edge of the Solar System.” Jet Propulsion Laboratory. California Institute of Technology. <http://stardust.jpl.nasa.gov/news/news101.html> (accessed on December 10, 2005).

“Crystalline Silica Exposure.” U.S. Occupational Safety and Health Administration. <http://www.osha.gov/Publications/osha3177.pdf> (accessed on December 10, 2005).

Scheer, James F. “Silica: Health and Beauty from Nature.” *Better Nutrition* (December 1997): 38–43. Also available online at http://www.findarticles.com/p/articles/mi_moFKA/is_n12_v59/ai_20086185 (accessed on December 10, 2005).

“Silica Gel.” J. T. Baker. <http://www.jtbaker.com/msds/englishhtml/s1610.htm> (accessed on December 10, 2005).

“Silica—Silicon Dioxide.” Azom.com. <http://www.azom.com/details.asp?ArticleID=1114> (accessed on December 10, 2005).

“What Is Silica Gel and Why Do I Find Packets of It in Everything I Buy?” How Stuff Works.
<http://science.howstuffworks.com/question206.htm> (accessed on December 10, 2005).

See Also Calcium Silicate; Sodium Silicate

OTHER NAMES:
Silver(I) iodide

FORMULA:
 AgI

ELEMENTS:
Silver, iodine

COMPOUND TYPE:
Binary salt (inorganic)

STATE:
Solid

MOLECULAR WEIGHT:
234.77 g/mol

MELTING POINT:
558°C (1036°F)

BOILING POINT:
1506°C (2743°F)

SOLUBILITY:
Insoluble in water and organic solvents; soluble in solutions of sodium chloride, potassium chloride, and ammonium hydroxide

KEY FACTS

Silver Iodide

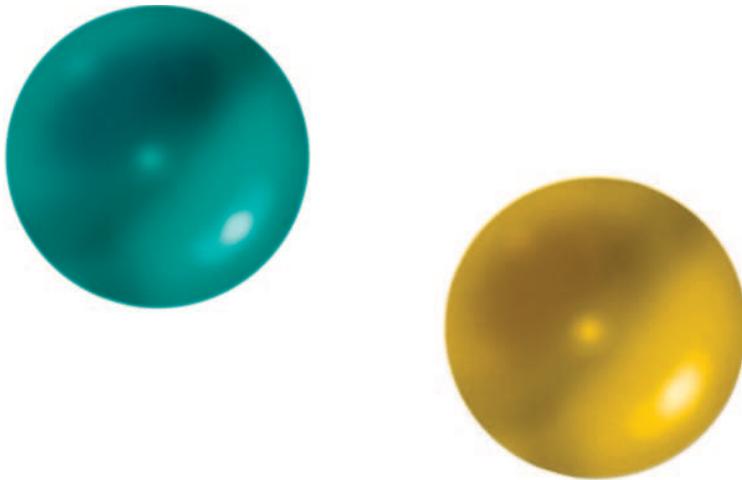
OVERVIEW

Silver iodide (SILL-ver EYE-oh-dide) is a light yellow crystalline or powdery material that darkens on exposure to light. The darkening occurs because silver ions (Ag^+ ; silver atoms with a positive charge) are converted to neutral silver atoms (Ag^0) that are dark gray in color. Silver iodide is used primarily in photography and in cloud-seeding experiments.

HOW IT IS MADE

Silver iodide occurs naturally in the mineral odargyrite (also known as iodyrite), from which it can be extracted. The compound is extracted by adding concentrated hydriodic acid (HI) to the mineral, which dissolves out the silver iodide. The silver iodide can then be recovered by evaporating the solution and purifying the product. The compound can also be made more easily by reacting silver nitrate (AgNO_3) with sodium or potassium iodide (NaI or KI) under a ruby red light. The colored light is used to prevent oxidation of the

Silver iodide. Turquoise atom is silver and yellow atom is iodine. PUBLISHERS RESOURCE GROUP



silver ion to silver metal during the preparation: $\text{AgNO}_3 + \text{NaI} \rightarrow \text{AgI} + \text{NaNO}_3$.

COMMON USES AND POTENTIAL HAZARDS

The primary use of silver iodide is in photography. The compound was first used for this purpose in the early nineteenth century by the French experimenter Louis Jacques Mandé Daguerre (1787–1851). Daguerre first covered a sheet of copper metal with a thin layer of silver. He then exposed the plate to iodine vapors, converting the silver metal to silver iodide. The photograph was taken by exposing the plate to light, which converted colorless silver ions in silver iodide back to grayish silver atoms. After exposure, the plate was treated with magnesium vapor, which adhered to the parts of the plate that had been exposed to light. The unreacted silver iodide was then washed off the plate, revealing an image on the plate. The fundamental principle invented by Daguerre is still used today in taking photographs with photograph film containing silver iodide.

The other major use for silver iodide is in cloud seeding. Cloud seeding is the process by which some foreign material—usually silver iodide or dry ice (solid carbon dioxide)—is dropped into a rain cloud. The crystals of silver iodide or carbon dioxide provide nuclei—tiny cores—on which water can condense to form water droplets. The process of cloud seeding was first developed in the 1940s by American chemist Vincent Schaefer (1927–1993). Schaefer used crushed

Interesting Facts

- The United States government spent about \$19 million per year in the 1970s for cloud seeding experiments. By the late 1990s, that amount had been

reduced to about \$500,000 as doubts were raised about the effectiveness of the procedure.

dry ice dropped from an airplane into a cloud in his first experiments. His initial experiments proved to be successful, with rain or snowstorms resulting from this seeding.

For three decades, researchers attempted to modify and improve Schaefer's techniques so as to be able to use cloud seeding to produce rain on demand. One step in that direction was the effort by General Electric scientist Bernard Vonnegut (1914-1997) to use silver iodide crystals rather than dry ice for seeding. Although these experiments produced moderately successful results, many scientists eventually lost confidence in the process of cloud seeding as a reliable method for producing rain. By the early twenty-first century, the process was being used on only a small scale in isolated areas in efforts to produce rain. The problem is that assessing the success of cloud seeding is very difficult. Because the process requires that clouds are already in an area, it is difficult to determine whether rain or snow falls as a result of cloud seeding or as a result of natural processes. In 2003, the National Academy of Sciences reached the conclusion that no reliable scientific evidence exists to suggest that cloud seeding produces more rain or snow than would occur naturally. The American Meteorological Society has taken a somewhat more optimistic view, suggesting that cloud seeding may increase precipitation by up to 10 percent.

Some concerns have been expressed about the environmental and health effects of using silver iodide for cloud seeding. However, only small amounts of silver iodide are released into the atmosphere. That which does fall to earth does not dissolve in water and so is unlikely to enter a community water supply. Tests have shown that the concen-

tration of silver iodide in rainwater is far below the 50 micrograms per liter that has been deemed safe by the U.S. Public Health Service. The primary health concern for workers who handle silver iodide is a condition known as argyreia, in which the skin is stained a bluish black color by the compound.

FOR FURTHER INFORMATION

“Chemistry of Photography.” Cheresources.

<http://www.cheresources.com/photochem.shtml> (accessed on November 3, 2005).

Fukuta, Norihiko. “Cloud Seeding Clears the Air.” *Physics in Action* (May 1998). Also available online at <http://physicsweb.org/articles/world/11/5/3/1> (accessed on November 3, 2005).

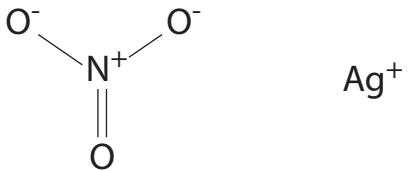
“Silver Iodide.” ESPI Metals.

<http://www.espmetals.com/msds/silveriodide.pdf> (accessed on November 3, 2005).

“Silver Iodide.” H&S Chemical.

<http://www.hschem.com/msds/silverIodide.htm> (accessed on November 3, 2005).

Zertuche, Casey. “Every Cloud Has a Silver Lining.” The Daily Texan (April 12, 2004). Also available online at http://www.dailytexanononline.com/media/paper410/news/2004/04/12/Focus/Every.Cloud.Has.A.Silver.Lining_657354.shtml?norewrite&sourcedomain=www.dailytexanononline.com (accessed on November 3, 2005).



OTHER NAMES:
Silver(I) nitrate; lunar caustic
FORMULA:
AgNO_3
ELEMENTS:
Silver, nitrogen, oxygen
COMPOUND TYPE:
Salt (inorganic)
STATE:
Solid
MOLECULAR WEIGHT:
169.87 g/mol
MELTING POINT:
212°C (414°F)
BOILING POINT:
440°C (824°F); decomposes
SOLUBILITY:
Soluble in water, glycerol, and hot ethyl alcohol; moderately soluble in acetone

KEY FACTS

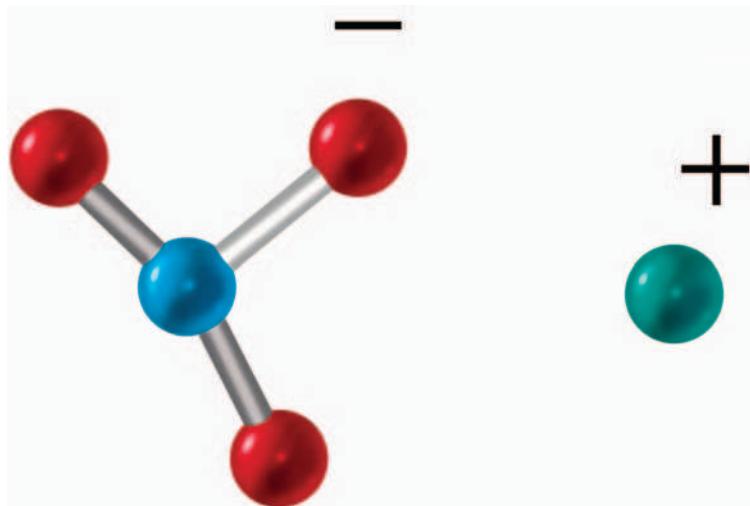
Silver Nitrate

OVERVIEW

Silver nitrate (SILL-ver NYE-trate) is a colorless to transparent to white crystalline solid with no odor and a bitter metallic taste. In pure form, the compound is not affected by light, but trace amounts of organic impurities may catalyze the conversion of silver ions (Ag^+ ; silver atoms with a positive charge) to grayish neutral silver atoms (Ag^0) that give the salt a grayish tint. Silver nitrate is the most widely used of all silver compounds, finding application in the synthesis of other silver compounds, as a catalyst in certain industrial chemical reactions, as an antiseptic and germicide, and in photographic processes.

The therapeutic effects of silver compounds, including silver nitrate, have been known for many centuries. Both the Greeks and Romans, for example, used aqueous solutions of silver nitrate to treat wounds and cuts. In 1881, the German physician Carl Crede (1819–1892) developed the practice of applying a 2 percent solution of silver nitrate to the eyes of newborn babies to prevent gonorrheal

Silver nitrate. Red atoms are oxygen; blue atom is nitrogen; and turquoise atom is silver. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP



ophthalmia, a bacterial infection of the eyes that may result in blindness in a child.

The use of silver nitrate in printing and photography dates to discoveries made in the 1720s by the German chemist Johann Schulze (1687–1744). Schulze found that a mixture of silver, nitric acid, and chalk turns purple or black when exposed to light. That simple discovery formed the basis of the modern science of photography. In 1802, Thomas Wedgwood (1771–1805), son of the founder of the famous Wedgwood pottery company, used silver nitrate to make temporary negative prints on paper, producing shades of gray as well as pure black and white. In 1835, the British mathematician William Henry Fox Talbot (1800–1877) made the first permanent paper negative from paper coated with silver nitrate and common table salt (sodium chloride).

Two years later, the French inventor Louis Jacques Mandé Daguerre (1787–1851) coated a copper plate with silver and washed it with nitric acid to create a plate on which he made the first Daguerreotype, an early form of photography. By the 1850s, Daguerre's technique was in wide use, with the substitution of glass plates coated with silver nitrate to obtain images produced by exposing the plates to light. By the end of the nineteenth century, emulsions of silver nitrate in celluloid were being used for making photographs, a process that was modified in the 1930s by the substitution of

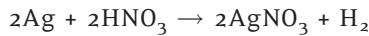
Interesting Facts

- Film speed numbers, such as ISO 100 or ISO 4000, correspond to the size of silver nitrate particles on the film. The finer the grain, the more detailed the image quality.
- Silver nitrate is not toxic to humans or other mammals. But it is toxic to fish and other aquatic organisms. For this reason, it should not be discarded in lakes and rivers.

cellulose acetate for celluloid. Silver nitrate remains an important component of the photographic process today.

HOW IT IS MADE

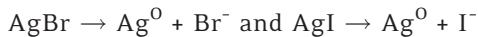
Silver nitrate is made by dissolving metallic silver in weak nitric acid.



The solution is then evaporated to recover the crystalline silver nitrate, which is heated to remove impurities, dissolved in water, and re-purified.

COMMON USES AND POTENTIAL HAZARDS

The primary use of silver nitrate is in the production of other silver salts used in the production of photographic film. Compounds such as silver bromide (AgBr) and silver iodide (AgI) decompose when exposed to light, forming free silver:



While silver bromide and silver iodide are nearly colorless, the free silver atoms formed in this reaction are black. Thus, portions of a photographic film that have been exposed to light turn black where free silver has formed.

In addition to its use in treating the eyes of newborn babies, silver nitrate has a number of other medical applications. It is used as a germicidal wall spray in medical facilities, as a topical (on the skin) anti-infective agent, for the cauterization of wounds, and as a general antiseptic. Cauterization is the process by which the skin surrounding a wound

Words to Know

ANTISEPTIC A chemical that prevents the growth of bacteria and viruses that cause disease.

AQUEOUS Referring to solution that consists of some material dissolved in water.

CATALYST A material that increases the rate of a chemical reaction without

undergoing any change in its own chemical structure.

EMULSION A temporary mixture of two liquids that normally do not dissolve in each other.

THERAPEUTIC DRUG A compound that has healing properties.

is burned in order to seal the wound. Other uses of silver nitrate include:

- The silvering of mirrors, a process by which a thin coating of silver metal is attached to the back of a piece of glass to form a mirror;
- As a catalyst in the manufacture of ethylene oxide, an important raw material in the production of plastics;
- In the silver plating of metals and plastics;
- In the manufacture of indelible printing inks;
- For hair dye; and
- As a flower preservative, a process that involves adding a small amount of silver nitrate solution to the water in which flowers are stored.

Silver nitrate is an irritant that can cause inflammation and burning of the skin, eyes, and respiratory system. It is also toxic by ingestion. In sufficient amounts, silver nitrate can cause severe damage to the respiratory tract and lungs, blindness, and even death. Solutions of silver nitrate can stain the skin dark purple or black, although such stains do not necessarily indicate that serious damage has occurred.

FOR FURTHER INFORMATION

Dunn, Peter M. "Dr Carl Credé (1819–1892) and the Prevention of Ophthalmia Neonatorum." *Archives of Diseases in Childhood* (September 2000): F158–F159. Also available online at <http://fn.bmjjournals.com/cgi/content/full/83/2/F158> (accessed on November 5, 2005).

Grimm, Tom, and Michele Grimm. *The Basic Book of Photography*. New York: Plume Books, 2003.

Patnaik, Pradyot. "Silver Nitrate." *Handbook of Inorganic Chemicals*. New York: McGraw Hill, 2003, 841 842.

"Silver Nitrate." New Jersey Department of Health and Senior Services.<http://nj.gov/health/eoh/rtkweb/1672.pdf> (accessed on November 5, 2005).

See Also Cellulose Nitrate; Silver Iodide



OTHER NAMES:	Silver oxide; argentous oxide
FORMULA:	Ag_2O
ELEMENTS:	Silver, oxygen
COMPOUND TYPE:	Metallic oxide (inorganic)
STATE:	Solid
MOLECULAR WEIGHT:	231.74 g/mol
MELTING POINT:	Decomposes at about 200°C (400°F)
BOILING POINT:	Not applicable
SOLUBILITY:	Slightly soluble in water; insoluble in ethyl alcohol

KEY FACTS

Silver(I) Oxide

OVERVIEW

Silver(I) oxide (SILL-ver one OK-side) is an odorless dark brown or black powder with a metallic taste. It is used primarily for polishing glass, the purification of water, and coloring glass.

HOW IT IS MADE

Silver(I) oxide is made by reacting silver nitrate (AgNO_3) with sodium or potassium hydroxide (NaOH or KOH). For example:

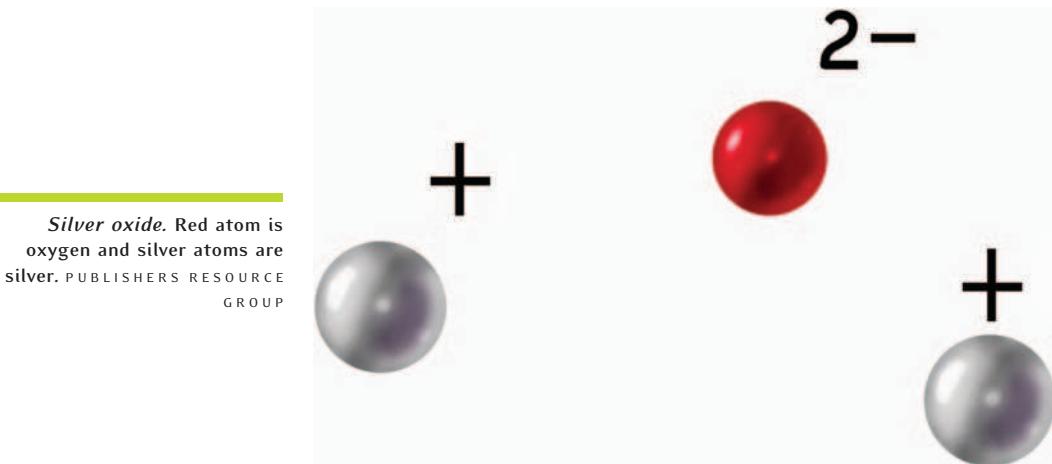


The silver(I) oxide settles out as a precipitate that can then be washed and purified.

COMMON USES AND POTENTIAL HAZARDS

Silver(I) oxide finds limited commercial and industrial application. It is used as an ingredient in the manufacture of

SILVER(I) OXIDE



glass to give a yellowish caste to the glass. It is also a component of mixtures used to polish glass, including the glass used in optical lenses. Silver(I) oxide is also used as a catalyst in certain industrial operations and in some water purification systems.

Silver(I) oxide is a skin, eye, and respiratory irritant that may cause coughing, wheezing, shortness of breath, and pulmonary edema (accumulation of fluid in the lungs). It can also cause burning of the eyes and skin. Ingestion can produce burning of the gastrointestinal tract accompanied by nausea, vomiting, and abdominal pain. Long-term exposure to silver(I) oxide can cause argyria, a bluish-gray discoloration of the skin, eyes, and mucous membranes (the soft tissues lining the breathing and digestive passages).

Words to Know

AQUEOUS Referring to a solution that consists of some material dissolved in water.

PRECIPITATE A solid material that settles out of a solution, often as the result of a chemical reaction.

FOR FURTHER INFORMATION

“Material Safety Data Sheet.” IC Controls.

<http://www.iccontrols.com/files/a1100122.pdf> (accessed on November 5, 2005).

“Silver(I) Oxide.” Patnaik, Pradyot. *Handbook of Inorganic Chemicals*. New York: McGraw Hill, 2003, 842 843.

“Silver Oxide.” Chemoo07.

http://www.chemoo07.com/specification_d/chemicals/supplier/cas/Silver%20oxide.asp (accessed on November 5, 2005).



OTHER NAMES:	Argentous sulfide
FORMULA:	Ag_2S
ELEMENTS:	Silver, sulfur
COMPOUND TYPE:	Binary salt (inorganic)
STATE:	Solid
MOLECULAR WEIGHT:	247.80 g/mol
MELTING POINT:	925°C (1520°F)
BOILING POINT:	Not applicable; decomposes
SOLUBILITY:	Insoluble in water; soluble in nitric acid, sodium cyanide (NaCN), and potassium cyanide (KCN)

KEY FACTS

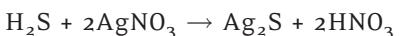
Silver(I) Sulfide

OVERVIEW

Silver(I) sulfide (SILL-ver one SUL-fide) is a grayish-black heavy powder. Most people are familiar with the compound as tarnish, the black coating that covers silver tableware and jewelry when they are exposed to the air.

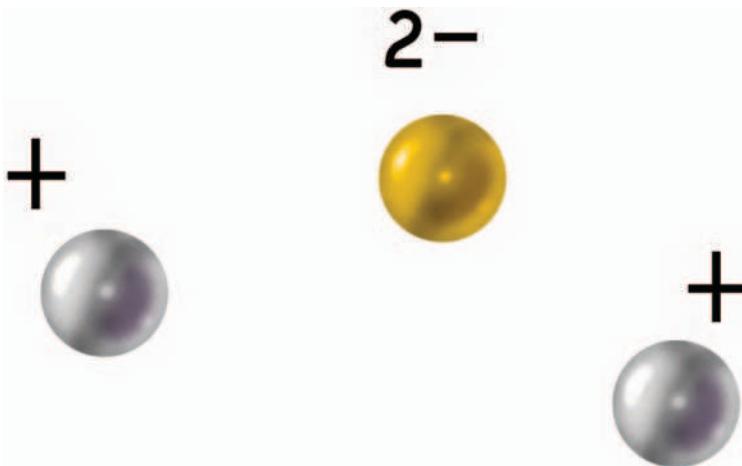
HOW IT IS MADE

Silver(I) sulfide occurs naturally as the minerals acanthite and argentite, from which they can be extracted by grinding, crushing, and washing the mineral ore. The compound can also be prepared synthetically by passing hydrogen sulfide (H_2S) gas through an aqueous solution of silver nitrate.



The silver(I) sulfide precipitates out of solution and can be filtered and purified by washing with hot water.

Silver sulfide. Silver atoms are silver and yellow atom is sulfur.
PUBLISHERS RESOURCE GROUP



COMMON USES AND POTENTIAL HAZARDS

Naturally occurring silver(I) sulfide can be used as a source of silver metal. The sulfide is roasted in air, converting silver(I) sulfide to silver(I) sulfate. The sulfate can then be treated chemically to obtain silver metal. The process finds little commercial application since other, more economically efficient, sources of silver are available.

The primary use of silver(I) sulfide is in the production of glazes for ceramics. The compound gives a glassy or metallic sheen to the glaze. It is also used in the process known as niello, first used by the early Greeks and Romans to produce a black, metallic inlay on the surface of pottery.

Silver(I) sulfide is an irritant to the skin, eyes, and respiratory system. No long-term effects of exposure to the compound have been reported. When absorbed through cuts in the skin or ingested it may produce a condition known as argyreia, which causes skin and mucous membranes to develop a bluish-black color.

FOR FURTHER INFORMATION

Krampf, Robert. "Cleaning the Silver." Edgerton Explorit Center.
<http://www.edgerton.org/kidscorner/silver.html> (accessed on November 5, 2005).

"Material Safety Data Sheet." ESPI Metals.
<http://www.espimetals.com/msds/s/silversulfide.pdf> (accessed on November 5, 2005).

Interesting Facts

- Tarnish is caused by a chemical reaction between silver in tableware and sulfur compounds that occur in eggs. Tarnish can be removed by soaking tableware in a solution of warm baking soda in water in a pan lined with aluminum foil. A chemical reaction occurs in which aluminum replaces silver, forming aluminum sulfide (Al_2S_3) and free silver:

$$2\text{Al} + 3\text{Ag}_2\text{S} \rightarrow 6\text{Ag} + \text{Al}_2\text{S}_3$$
The reaction occurs only if the silver tableware is in physical contact with the aluminum foil because an electric current must flow between the two metals.
- The silver present in electric contacts exposed to hydrogen sulfide gas may begin to grow long filaments known as *silver whiskers*. These whiskers have been known to grow as long as 8 centimeters (3 inches) in length and can cause catastrophic failures in the electrical contacts.

"Mineral Acanthite/Argentite, The." Amethyst Galleries.
<http://mineral.galleries.com/minerals/sulfides/acanthit/acanthit.htm> (accessed on November 5, 2005).

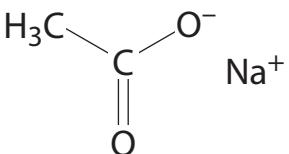
Patnaik, Pradyot. *Handbook of Inorganic Chemicals*. New York: McGraw Hill, 2003, 845.

Words to Know

AQUEOUS Referring to a solution that consists of some material dissolved in water.

PRECIPITATE A solid material that settles out of a solution, often as the result of a chemical reaction.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.



FORMULA:	$\text{NaC}_2\text{H}_3\text{O}_2$
ELEMENTS:	Sodium, carbon, hydrogen, oxygen
COMPOUND TYPE:	Salt (inorganic)
STATE:	Solid
MOLECULAR WEIGHT:	82.03 g/mol
MELTING POINT:	328.2°C (622.8°F)
BOILING POINT:	Not applicable; decomposes
SOLUBILITY:	Very soluble in water; soluble in ether; slightly soluble in ethyl alcohol

KEY FACTS

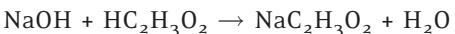
Sodium Acetate

OVERVIEW

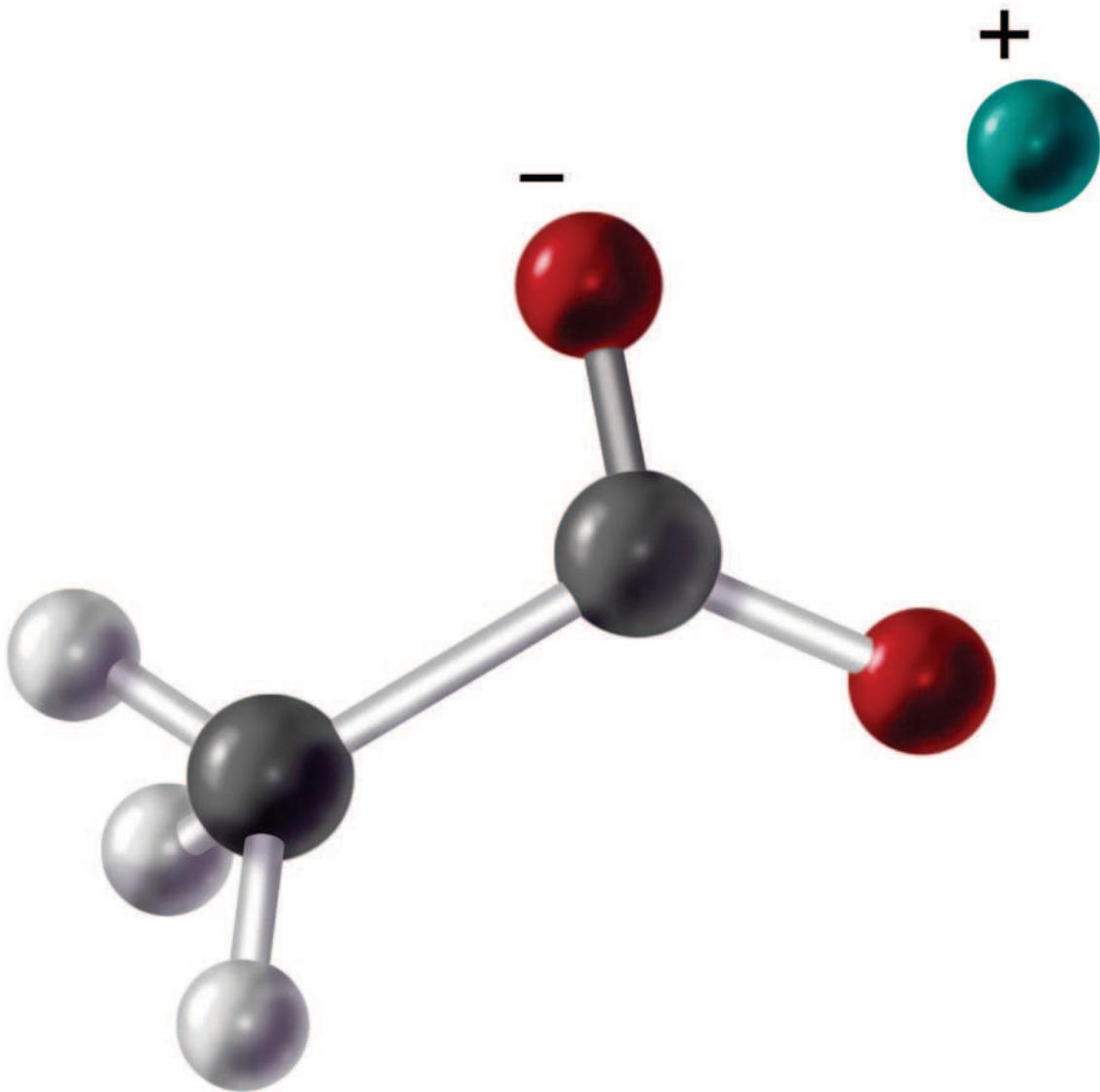
Sodium acetate (SO-dee-um ASS-uh-tate) is a colorless, odorless crystalline solid that often occurs as the trihydrate: $\text{NaC}_2\text{H}_3\text{O}_2 \cdot 3\text{H}_2\text{O}$. A hydrate is a chemical compound formed when one or more molecules of water is physically added to the molecule of some other substance. Sodium acetate trihydrate has three molecules of water of hydration for every $\text{NaC}_2\text{H}_3\text{O}_2$ unit. Anhydrous sodium acetate readily converts to the trihydrate because it is very hygroscopic. A hygroscopic compound is one that readily absorbs moisture from the air.

HOW IT IS MADE

Sodium acetate is prepared by reacting either sodium hydroxide (NaOH) or sodium carbonate (Na_2CO_3) with acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$). With sodium hydroxide, for example, the reaction is:



The sodium acetate forms as the trihydrate, which is then recovered by evaporating the reacting solution. The



Sodium acetate trihydrate.

Red atoms are oxygen;
hydrogen atoms are white;
black atoms are carbon;
turquoise atom is sodium. Gray
stick indicates double bond.

PUBLISHERS RESOURCE GROUP

anhydrous salt can be prepared by heating the hydrate to drive off the water of hydration.

COMMON USES AND POTENTIAL HAZARDS

Sodium acetate is used in a number of industries. In the textile industry, it is used as a mordant in the dyeing of fabrics and to stabilize and improve the finish of fabrics. The

Interesting Facts

- In 1995, officials in Michigan tested sodium acetate trihydrate as a deicing salt for airport runways. Experiments showed, however, that the anhydrous salt worked better than the trihydrate for this purpose.
- Sodium acetate has also been used as a deicer in parking garages. The compound is preferred to sodium chloride, widely used as a deicing compound, because sodium chloride corrodes steel rods buried in concrete and sodium acetate does not.
- Heat packs sometimes contain a solution of sodium acetate trihydrate cooled below its freezing point. When the pack is activated, the compound freezes very rapidly, releasing the heat that had been stored in its super-cooled phase.

cosmetics industry uses sodium acetate as a buffering agent in a variety of personal care products. A buffering agent is a substance that maintains the acidity of a product within a certain desired range. Sodium acetate is also used by food producers for the same reason, assuring that a variety of foods have an acidity sufficient to protect the food from decaying, but not so acidic as to have an unpleasant taste. Some other applications of sodium acetate include:

- In heat packs to relieve stiffness and pain, to keep hands and feet warm, and to warm baby bottles;
- In the production of soaps, where it reacts with strong bases to reduce the harshness of the final product;
- In dialysis machines, used for people whose kidneys are not working properly, to provide the sodium ions (Na^+) to maintain proper electrolyte balance in the body;
- As a diuretic, a drug used to promote urination;
- As an expectorant in drugs used to promote coughing to help bring up mucous;

Words to Know

ANHYDROUS Lacking water of hydration.

MORDANT A substance used in dyeing and printing that reacts chemically with both a

dye and the material being dyed to help hold the dye permanently to the material.

- As a veterinary treatment for bovine ketosis, a condition caused by low blood sugar in cows that results in a wasting or weakening of the animal;
- As a buffer in the developing of photographs;
- In the tanning of hides to obtain a more even and more rapid absorption of the tanning material; and
- In the purification of glucose.

Sodium acetate is a mild irritant to the skin, eyes, and respiratory system. If inhaled, it may cause inflammation of the throat and lungs. At the level it appears in most household products, it presents a very low hazard to the average person.

FOR FURTHER INFORMATION

“How Do Sodium Acetate Heat Pads Work?” How Stuff Works.

<http://home.howstuffworks.com/question290.htm> (accessed on November 5, 2005).

“Material Safety Data Sheet: Sodium Acetate Trihydrate.” Iowa State University, Department of Chemistry.

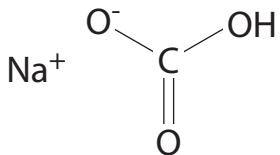
http://avogadro.chem.iastate.edu/MSDS/NaOAc_3H2O.htm (accessed on November 5, 2005).

“Sodium Acetate, Anhydrous.” Cornell Material Safety Data Sheets.

<http://msds.ehs.cornell.edu/msds/msdsdod/a73/m36008.htm#Section3> (accessed on November 5, 2005).

“Sodium Acetate Anhydrous Physical Properties.” Jarchem Industries.

http://www.jarchem.com/sodium_acetate_anhydrous.htm (accessed on November 5, 2005).

**OTHER NAMES:**

Bicarbonate of soda;
baking soda

FORMULA:

NaHCO_3

ELEMENTS:

Sodium, hydrogen,
carbon, oxygen

COMPOUND TYPE:

Acid salt (inorganic)

STATE:

Solid

MOLECULAR WEIGHT:

84.01 g/mol

MELTING POINT:

about 50°C (120°F);
decomposes

BOILING POINT:

Not applicable;
decomposes

SOLUBILITY:

Soluble in water;
insoluble in ethyl
alcohol

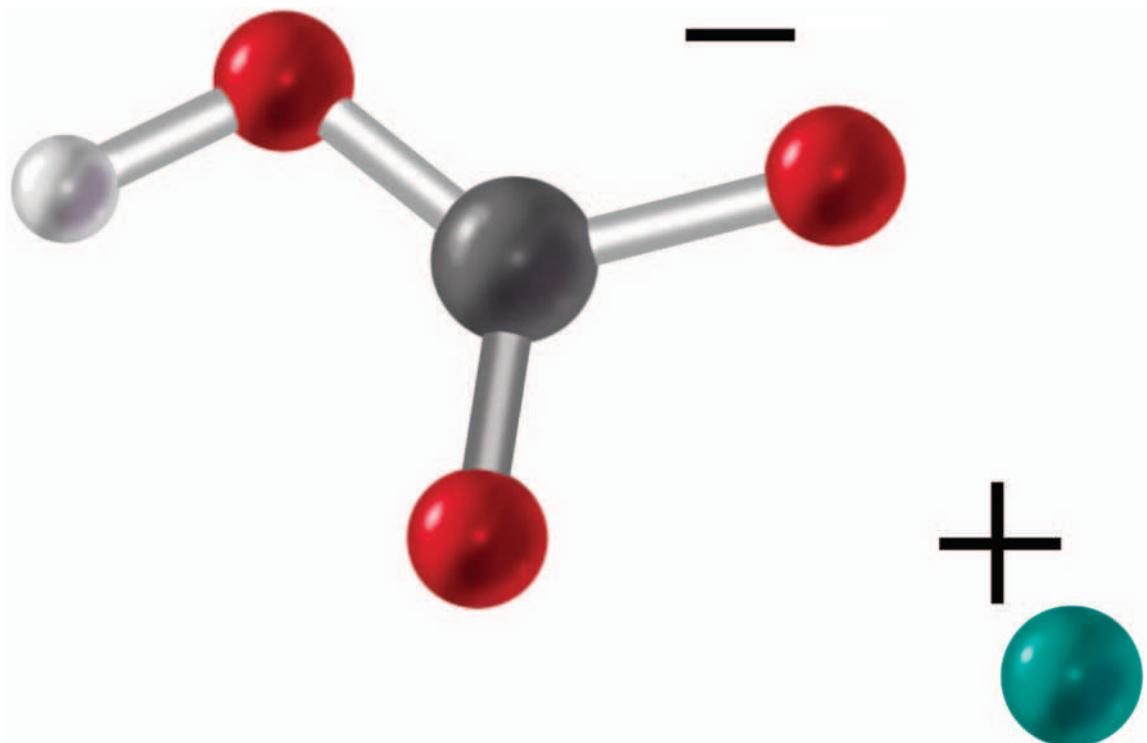
KEY FACTS**OVERVIEW**

Sodium bicarbonate (SO-dee-um bye-KAR-bun-ate) is a white, odorless, crystalline solid or powder that is stable in dry air, but that slowly decomposes in moist air to form sodium carbonate. The compound's primary uses are as an additive in human and animal food products.

Sodium bicarbonate has been used by humans for thousands of years. Ancient Egyptian documents mention the use of a sodium bicarbonate and sodium chloride solution in the mummification of the dead. For centuries, people around the world have used sodium bicarbonate as a leavening agent for baking. A leavening agent is a substance that causes dough or batter to rise. Sodium bicarbonate produces this effect because, when heated or dissolved in water, it breaks down to produce carbon dioxide (CO_2) gas:



SODIUM BICARBONATE



Sodium bicarbonate. Red atoms are oxygen; white atom is hydrogen; black atom is carbon; and turquoise atom is sodium. Gray stick indicates double bond. PUBLISHERS RESOURCE GROUP

Since all the compounds present in this reaction are safe for human consumption, sodium bicarbonate makes an ideal leavening agent.

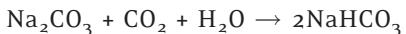
Commercial production of sodium bicarbonate as baking soda dates to the late 1700s. In 1846, Connecticut physician Austin Church (1799–?) and John Dwight (1819–?) of Dedham, Massachusetts, founded a company to make and sell sodium bicarbonate. They started their company in the kitchen of Dwight's home, making the product by hand and packing it in paper bags for sale to neighbors. The Church-Dwight operation grew over the years to become the largest producer of household baking soda, now sold under the name of Arm & Hammer® baking soda. The company still produces 90 percent of all the baking soda used for household purposes in the United States. Consumers use the product for cooking, cleaning, and deodorizing homes.

Interesting Facts

- Sodium bicarbonate is a very effective cleaning agent for certain materials. In the 1980s, restorers used an aqueous solution of the compound to clean the Statue of Liberty.

HOW IT IS MADE

Sodium bicarbonate is made commercially by one of two methods. In the first method, carbon dioxide gas is passed through an aqueous solution of sodium carbonate (Na_2CO_3):



Since the bicarbonate is less soluble than the carbonate, it precipitates out of solution and can be removed by filtration.

Sodium bicarbonate is also obtained as a byproduct of the Solvay process. The Solvay process was invented in the late 1850s by Belgian chemist Ernest Solvay (1838–1922) primarily as a way of making sodium carbonate. Sodium carbonate had long been a very important industrial chemical for which no relatively inexpensive method of preparation existed. Solvay developed a procedure by which sodium chloride is treated with carbon dioxide and ammonia, resulting in the formation of sodium bicarbonate and ammonium bicarbonate. The sodium bicarbonate is then heated to obtain sodium carbonate. Although sodium carbonate is the desired product in this reaction, sodium bicarbonate can also be obtained by deleting the final step by which it is converted into sodium carbonate.

COMMON USES AND POTENTIAL HAZARDS

An estimated 560,000 metric tons (615,000 short tons) of sodium bicarbonate were consumed in the United States in 2003. About one-third of that amount was used by the food products industry, primarily in the manufacture of baking soda (pure sodium bicarbonate) and baking powder (a mixture of sodium bicarbonate and at least one other compound).

Baking powder differs from baking soda in that it includes an acidic compound that reacts with sodium bicarbonate to produce carbon dioxide. One of the most common compounds mixed with sodium bicarbonate in baking powder is tartaric acid ($\text{HOOC}(\text{CHOH})_2\text{COOH}$), or its salt, potassium bitartrate ($\text{HOOC}(\text{CHOH})_2\text{COOK}$). Baking powder is a more efficient leavening agent in baking than is baking soda by itself. Baking soda is also used as an additive in foods and drinks to provide effervescence (a bubbling, fizzing, or sparkling effect) or to maintain an acidic environment in the food. The acidity provides a sharp taste and helps to preserve a food.

The second largest use of sodium bicarbonate is as an additive in animal feed. As with human foods, it maintains the proper acidity of an animal's feed, improving its ability to digest and absorb its food.

Sodium bicarbonate is also used in a number of pharmaceutical applications. For example, it is a common ingredient in antacids, products designed to relieve heartburn, acid indigestion, sour stomach, and other discomforts caused by overeating or improper foods. Some pharmaceuticals, such as Alka-Seltzer®, contain a combination of citric acid and sodium bicarbonate. The citric acid helps the sodium bicarbonate dissolve more quickly and produces more effervescence when the tablet is dissolved in water.

Sodium bicarbonate is also used in cleaning products on both a household and industrial level. Many householders use commercial baking soda, such as that sold by the Arm & Hammer company, to clean kitchen and bathroom appliances, such as sinks, stoves, and toilet bowls. Industries also use sodium bicarbonate filters to remove sulfur dioxide and other pollutants in flu gases released from factory smokestacks. The compound is also used in the treatment of wastewater to maintain proper acidity, remove certain odors (such as those of sulfur dioxide), and destroy bacteria. Some communities have used aqueous solutions of sodium bicarbonate sprayed at high pressure to remove graffiti; paint; soot and smoke residues; and mold from buildings, walls, and other public structures.

Some other applications of sodium bicarbonate include:

- As a component of fire extinguishers; when it comes into contact with an acid in the fire extinguisher, the sodium bicarbonate releases carbon dioxide and a flow of water under pressure to fight the fire;

Words to Know

AQUEOUS SOLUTION A solution that consists of some material dissolved in water.

PRECIPITATE A solid material that settles out of a solution, often as the result of a chemical reaction.

- As a blowing agent in the preparation of plastics; blowing agents are substances that produce large volumes of gas that convert a molten product into a foamy product;
- In the manufacture of other sodium compounds;
- For gold and platinum plating; and
- To prevent the growth of mold on timber.

Sodium bicarbonate is considered safe when handled or ingested in reasonable amounts. As with any chemical, however, excessive amounts of the compound can have harmful effects. When ingested in large amounts, sodium bicarbonate can cause stomach cramps, gas, upset stomach, vomiting, frequent urination, loss of appetite, and blood in the urine and stools.

FOR FURTHER INFORMATION

“Pure Baking Soda.” Arm & Hammer®.

<http://www.armhammer.com/> (accessed on November 8, 2005).

Snyder, C. H. *The Extraordinary Chemistry of Ordinary Things*, 4th ed. New York: John Wiley and Sons, 2002.

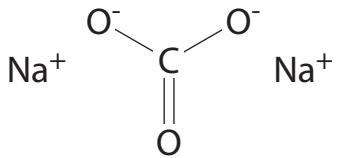
“Sodium Bicarbonate.” Chemical Land 21.

<http://www.chemicaland21.com/arokorhi/industrialchem/inorganic/SODIUM%20BICARBONATE.htm> (accessed on November 8, 2005).

“Sodium Bicarbonate.” DC Chemical Co., Ltd.

http://www.dcchem.co.kr/english/product/p_basic/p_basico2.htm (accessed on November 8, 2005).

See Also Carbon Dioxide; Citric Acid; Sodium Carbonate



OTHER NAMES:

Anhydrous salt; soda ash; Solvay soda; decahydrate; sal soda; washing soda (see Overview)

FORMULA:

Na_2CO_3

ELEMENTS:

Sodium, carbon, oxygen

COMPOUND TYPE:
Salt (inorganic)

STATE:
Solid

MOLECULAR WEIGHT:
105.99 g/mol

MELTING POINT:
858.1°C (1576°F)

BOILING POINT:
Not applicable;
begins to decompose
below melting point

SOLUBILITY:
Soluble in water;
insoluble in ethyl
alcohol

K E Y F A C T S

OVERVIEW

Sodium carbonate (SO-dee-um KAR-bun-ate) is an odorless white powder or crystalline solid with an alkaline taste. (Baking soda is another substance with an alkaline taste.) It is hygroscopic, meaning that it has a tendency to absorb moisture from the air. It also exists as the monohydrate ($\text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O}$) and as the decahydrate ($\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$), each with slightly different physical properties from those of the anhydrous salt. The anhydrous form of sodium carbonate is commonly known as soda ash, while the decahydrate is often called sal soda or washing soda. Sodium carbonate has long been one of the most important chemical compounds produced in the United States. Its primary use is in the manufacture of glass and other chemicals.

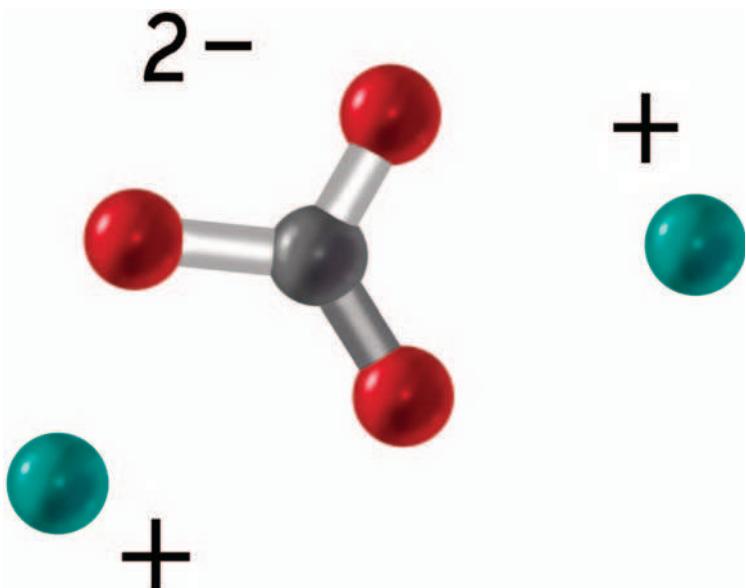
HOW IT IS MADE

Humans have known about and used sodium carbonate for thousands of years. The ancient Egyptians extracted the compound from a mineral known as natron found in dry lake

SODIUM CARBONATE

Sodium carbonate. Red atoms are oxygen; black atom is carbon; and turquoise atoms are sodium. Gray stick indicates double bond.

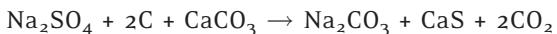
PUBLISHERS RESOURCE GROUP



bottoms. Natron is a combination of sodium carbonate and sodium bicarbonate. The Egyptians used sodium carbonate in the mummification of dead bodies. The compound dried out the bodies of the dead and prevented them from decaying. The technique was so effective that some mummified bodies over 3,000 years old are in as good a condition today as they were when the person died. Over the centuries, sodium carbonate was also produced by the combustion of organic matter, especially seaweed. This method of production accounts for the compound's common name of soda ash (ash containing sodium compounds).

The burning of dead plants does not produce very large quantities of sodium carbonate, so early chemists searched for synthetic methods of producing the increasingly important compound. The first breakthrough in that search occurred in 1791 when French chemist Nicolas Leblanc (1742-1806) invented a method for making sodium carbonate that became the industry standard for nearly a century. In the Leblanc method, sodium chloride (NaCl) is treated with sulfuric acid (H_2SO_4) to make sodium sulfate (Na_2SO_4) and hydrochloric acid (HCl). The sodium sulfate is then heated with charcoal (nearly pure carbon) and limestone ($CaCO_3$).

The product of this reaction is a dark ashy material that contains sodium carbonate, calcium sulfide, carbon dioxide, and other byproducts:



The sodium carbonate is extracted from this mixture and purified. The Leblanc invention was one of the great breakthroughs in the early years of chemical science. It made possible, among other things, the mass production of soap for the first time in human history.

As important as Leblanc's invention was, it suffered from one serious drawback: It required large amounts of energy. For this reason, chemists were always on the lookout for an alternative method for producing sodium carbonate that was less energy-intensive. That breakthrough came in 1861 when Belgian chemist Ernest Solvay (1838–1922) found a new way to make the important compound. Solvay found that treating sodium chloride with carbon dioxide and ammonia resulted in the formation of sodium bicarbonate and ammonium bicarbonate. Simply heating the sodium bicarbonate converts the bicarbonate to the carbonate. Like Leblanc's discovery, the Solvay process is regarded as one of the great accomplishments in the first century of industrial chemistry. By 1900, almost all of the sodium carbonate produced in the world was being made by the Solvay process.

That situation has changed. Today the most important source of sodium carbonate is natural minerals, such as thermonatrite (sodium carbonate monohydrate; $\text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O}$) or natron (or natrite; sodium carbonate decahydrate; $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$). These minerals are obtained from rocky deposits or from brines that are rich in the compound. Brine is water that is saturated with salts, such as sodium chloride, potassium chloride, and sodium carbonate. It is similar to, but saltier than, seawater.

COMMON USES AND POTENTIAL HAZARDS

About half of all the sodium carbonate produced in the United States is used to make glass products such as glass containers, flat glass, and fiber insulation. Glass is made by fusing (melting) a mixture of silica (silicon dioxide; SiO_2), sodium carbonate, limestone (calcium carbonate; CaCO_3), and other materials. The recipe differs considerably depending

Interesting Facts

- Over the last decade, consumption of sodium carbonate in the United States has been decreasing, largely because glass containers (made with sodium carbonate) are being replaced by plastic containers. However, the

industry has not gone into decline primarily because the demand for glass in developing countries continues to increase at a rate that matches the decline in glass bottle demand in the United States.

on the type of glass one wishes to make. For example, oxides of various metals such as iron(III) oxide and copper(II) oxide are added to provide a reddish tint to the glass.

The second highest use of sodium carbonate is in the production of other chemical compounds, followed by the compound's use in the production of soaps and detergents. Other applications of the compound include:

- The production of pulp and paper products;
- The removal of sulfur dioxide from flu gases in factories;
- In water purification and waste water treatment facilities;
- As a mordant in the dyeing of cloth;
- In the refining of petroleum;
- As a catalyst in the process by which coal is converted into a liquid fuel;
- For the bleaching of cotton and linen fabrics;
- As an emetic (a compound that induces vomiting); and
- In cosmetics and personal care products because of its ability to clean skin and help clear up rashes.

No serious health hazards have been associated with the use of sodium carbonate in any of its forms.

Words to Know

ANHYDROUS Lacking water of hydration.

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

MORDANT A substance used in dyeing and printing that reacts chemically with both a dye and the material being dyed to help hold the dye permanently to the material.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

WATER OF HYDRATION Water that has combined with a compound by some physical means.

FOR FURTHER INFORMATION

Kiefer, David M. "Soda Ash, Solvay Style." *Today's Chemist* (February 2002): 87–88+. Also available online at <http://pubs.acs.org/subscribe/journals/tcaw/11/i02/html/02chemchron.html> (accessed on November 8, 2005).

Lister, Ted, compiler. "Sodium Carbonate A Versatile Material." Royal Society of Chemistry. http://www.chemsoc.org/pdf/LearnNet/rsc/SodiumCarb_sel.pdf (accessed on November 8, 2005).

Monet, Jefferson. "An Overview of Mummification in Ancient Egypt." TourEgypt.net. <http://www.touregypt.net/featurestories/mummification.htm> (accessed on November 8, 2005).

"Soda Ash or Trona." Mineral Information Institute. <http://www.mii.org/Minerals/phototrona.html> (accessed on November 8, 2005).

"Sodium Carbonate." United Nations Environmental Programme. <http://www.inchem.org/documents/sids/sids/Naco.pdf> (accessed on November 8, 2005).

See Also Sodium Bicarbonate



OTHER NAMES:
Salt; table salt; common salt; rock salt
FORMULA:
NaCl
ELEMENTS:
Sodium, chlorine
COMPOUND TYPE:
Binary salt (inorganic)
STATE:
Solid
MOLECULAR WEIGHT:
58.44 g/mol
MELTING POINT:
800.7°C (1473°F)
BOILING POINT:
1465°C (2669°F)
SOLUBILITY:
Soluble in water and glycerol; very slightly soluble in ethyl alcohol and methyl alcohol

KEY FACTS

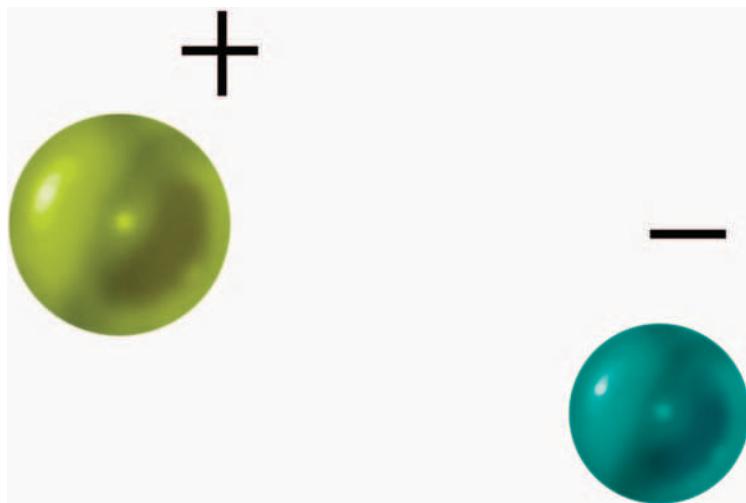
Sodium Chloride

OVERVIEW

Sodium chloride (SO-dee-um KLOR-ide) is a colorless to white powder or crystalline solid with no odor and a characteristic salty taste. It is slightly hygroscopic, meaning that it tends to absorb moisture from the air and become damp.

Salt is probably one of the best known and most widely used of all chemical compounds. Humans have been using salt as a preservative and to flavor foods since the beginning of recorded time. One of the earliest mentions of sodium chloride dates to 2,700 BCE in the Chinese book *Peng Tzao Kan Mu*, probably the first book on pharmacology ever written. Access to salt resources has often been a contentious issue among peoples, leading to battles and wars over its ownership. It has been considered at times to be so valuable that it was used as a form of money. Today, sodium chloride has a host of applications beyond its use as a food additive.

Sodium chloride. Turquoise atom is sodium and green atom is chloride. PUBLISHERS RESOURCE GROUP



HOW IT IS MADE

Sodium chloride occurs naturally as the mineral halite and abundantly in the oceans, where it is found in seawater at an average concentration of about 2.6 percent. There are several methods for harvesting salt, some of which date to ancient times. The earliest known method of production is also the simplest: evaporation of seawater by the Sun. In this method, seawater is collected in large, shallow ponds and allowed to evaporate. The salts dissolved in the water crystallize on the bottom of the ponds and can be scraped off and the individual compounds present—including sodium chloride—separated from each other.

This method works best in hot, arid parts of the world. In cooler, moister regions, seawater must be collected in large containers that can be heated artificially. In many cases, the seawater is heated under reduced pressure to allow it to boil at a lower temperature and save heating costs. Again, crystals of sodium chloride (and other dissolved salts) form as the water boils away.

Perhaps the most important source of sodium chloride is salt mines, large underground reserves of sodium chloride left behind when ancient seas dried up and were buried by the accumulation of rocks and soil. Salt mines are found in many parts of the world, especially Russia, Germany, the United Kingdom, India, France, Mexico, Canada, and the United States. These mines often span many kilometers and

extend hundreds of meters deep. One of the most famous salt mines in the United States is located under the city of Detroit. It contains more than 80 kilometers (50 miles) of underground roads built to remove blocks of sodium chloride, some as wide as four-lane highways. The Detroit mine ceased production in 1983 when lower salt prices and resulting lower profits were no longer able to sustain the costs of extraction. The Detroit mine was reopened in 1988, but only for the mining of road salt.

Two methods of mining are used to remove sodium chloride from underground sources. In the room-and-pillar method, shafts are dug into the deposit. Drilling and blasting are then used to break off pieces of sodium chloride, which are removed for processing. As mining progresses, large pillars of salt are left standing to support the empty chambers. Thus the name: room-and-pillar. A second method of salt removal is solution mining. A well is drilled into the ground and flooded to create a saturated solution of sodium chloride, a *brine* solution. The brine is then pumped to the surface and processed.

COMMON USES AND POTENTIAL HAZARDS

Probably the best known use of sodium chloride is as a food additive. The Salt Institute estimates that humans consume an average of 16 tons of salt during their lifetimes. Salt has long been used on foods to improve flavor and also as a preservative. Salted foods last longer than unsalted foods because salt inhibits the growth of microorganisms that cause decay.

Still, the most important use of sodium chloride by far is as a raw material in the production of other compounds. In 2004, 65 percent of all the sodium chloride consumed in the United States was used in the production of sodium hydroxide, sodium carbonate, hydrogen chloride, sodium metal, chlorine gas, and other chemical products. The next most important use of sodium chloride is in water conditioners. The compound is used in such devices because the sodium in sodium chloride will replace the calcium and magnesium in "hard" water (water in which it is hard to make suds). By softening water with sodium chloride, clothing and other materials can be cleaned more efficiently at lower cost. The Salt Institute claims that sodium chloride has more than 14,000 distinct uses. Some of the most important of those uses include:

Interesting Facts

- Salt was a key ingredient in the solution used by ancient Egyptians in the preparation of mummies.
- The expression “not worth his salt” had its origin in the ancient Greek slave trade, in which people were bought and sold for measures of salt.
- The English word *salary* comes from the Latin term *salarium argentum*, which refers to a special salt ration given to Roman soldiers.
- Salt has frequently been subject to heavy taxation; the very high tax on salt in France during the middle eighteenth century contributed to the rise of the French revolution.
- Salt mines that are no longer in use are sometimes used to store natural gas and petroleum.

- As a feed additive for livestock, poultry, and other domestic animals, to ensure that they receive the sodium and chlorine they need to remain healthy and grow normally;
- As a deicing product on roads and highways;
- In the manufacture of glazes used on ceramic products;
- For the curing of animal hides;
- In the dyeing and printing of fabrics;
- In the manufacture of soaps;
- As a herbicide, a chemical used to kill weeds; and
- As a fire extinguisher for certain types of fires (such as grease fires).

Given its widespread use, sodium chloride is obviously safe for consumption by most humans under normal conditions. As with any chemical compound, consumption of a large excess of sodium chloride can be harmful. The one health issue of greatest concern has to do with high blood pressure. Scientists have learned that the ingestion of large amounts of sodium can contribute to hypertension (high blood pressure), which in turn is associated with increased risk for heart attacks and stroke.

Words to Know

PHARMACOLOGY The study of compounds used as drugs.

The American Heart Association recommends that healthy American adults consume no more than 2,300 milligrams of sodium a day. That amount is equivalent to about a teaspoon of salt. For those considered at higher risk, people with high blood pressure, blacks, and middle-aged and older adults, the 2005 U.S. Department of Agriculture guidelines recommend no more than 1,500 milligrams of sodium per day. The problem with sodium chloride consumption is that most people have no idea how much salt they eat every day. Of course, they can keep track of the salt they add to the foods they prepare in their own homes. But most commercially prepared foods also have sodium chloride added to them. In some cases, the total amount of salt ingested from processed foods by the average American can be significant, easily exceeding the recommended daily average recommended by the American Heart Association. People can, therefore, be consuming dangerously high levels of sodium without being aware of that fact.

FOR FURTHER INFORMATION

“History of Salt.” The Salt Institute.

<http://www.saltinstitute.org/38.html> (accessed on November 8, 2005).

Kurlansky, Mark. *Salt: A World History*. New York: Walker, 2002.

“Salt.” History for Kids.

<http://www.historyforkids.org/learn/food/salt.htm> (accessed on November 8, 2005).

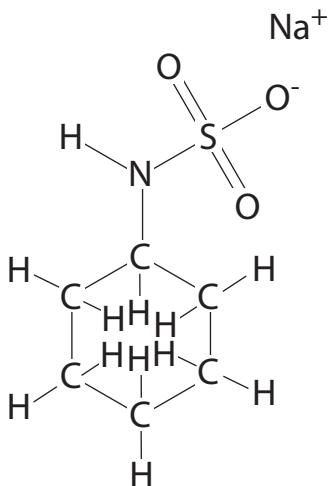
“Sodium Chloride.” J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/S3338.htm> (accessed on November 8, 2005).

“What You Always Wanted to Know about Salt.” The Salt Institute.

<http://www.saltinstitute.org/4.html> (accessed on November 8, 2005).

See Also Chlorine; Sodium Hydroxide



OTHER NAMES:

Monosodium cyclohexylsulfamic acid; sodium cyclohexane-sulfamate

FORMULA:
 $C_6H_{11}NHSO_3Na$

ELEMENTS:
Carbon, hydrogen, nitrogen, sulfur, oxygen, sodium

COMPOUND TYPE:
Organic salt

STATE:
Solid

MOLECULAR WEIGHT:
201.22 g/mol

MELTING POINT:
265°C (509°F)

BOILING POINT:
Not applicable;
decomposes

SOLUBILITY:
Soluble in water;
insoluble in most
organic solvents

K E Y F A C T S

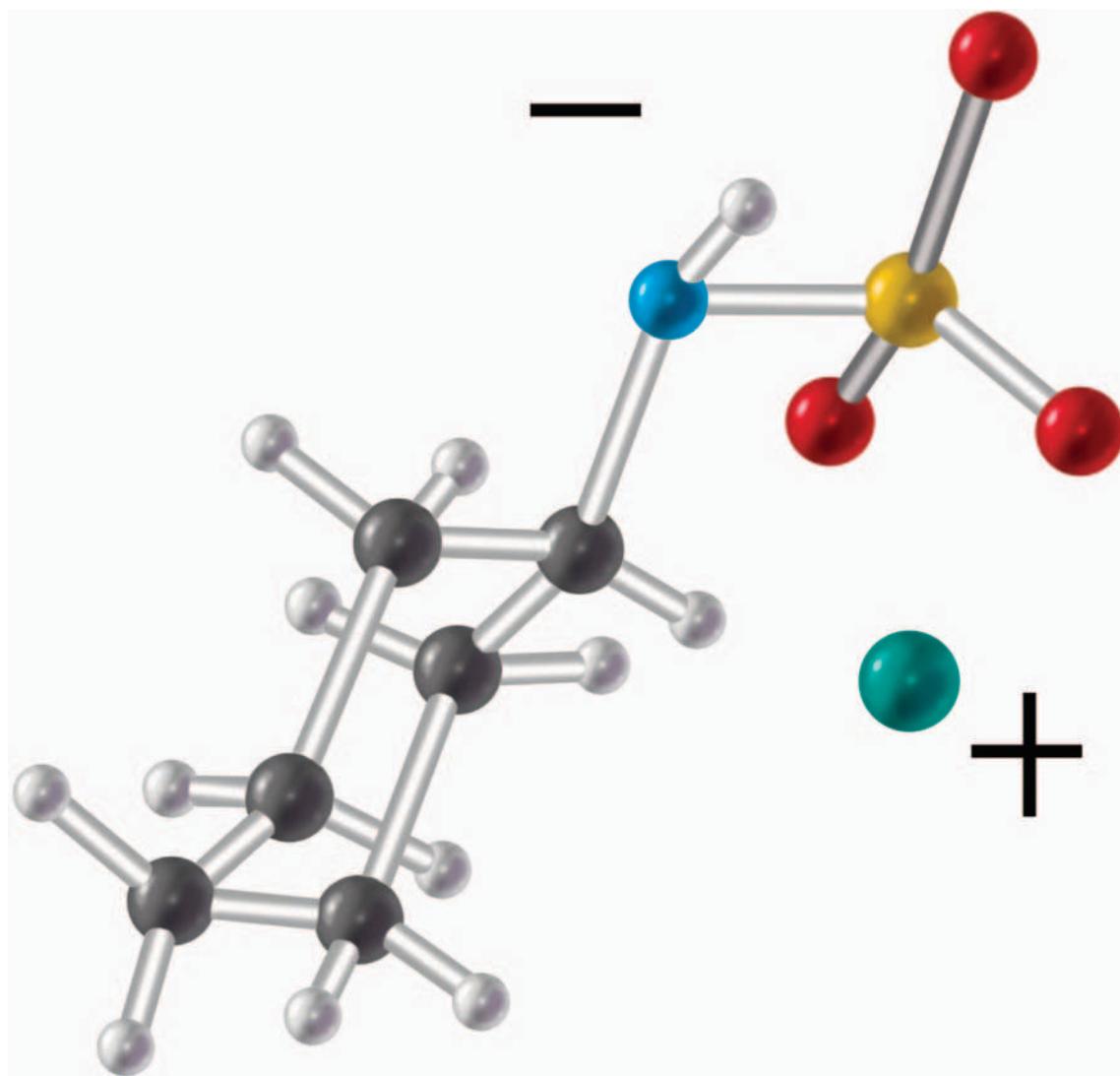
Sodium Cyclamate

OVERVIEW

Sodium cyclamate (SO-dee-um SYE-kla-mate) is a white, crystal solid or powder with almost no odor and a very sweet taste. Its sweetening power is about 30 times that of table sugar, the standard against which artificial sweeteners are measured. Because of its sweet flavor, sodium cyclamate is used as an artificial sweetener.

The cyclamate family of compounds was discovered in 1937 by Michael Sveda (1912-1999), then a graduate student at the University of Illinois. Sveda was working on the development of new drugs to treat fever. The story is that Sveda was smoking while he was working in the laboratory (a practice that would not be allowed today) and, at one point, he brushed some loose threads of tobacco from his lips. As he did so, he noticed a very sweet flavor on the cigarette. Curious as to the cause of the sweetness, Sveda looked more carefully into the compounds he was studying and eventually identified the source of the sweetness as a substance belonging to a class of compounds known as cyclamates. The cyclamates

SODIUM CYCLAMATE



Sodium cyclamate. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; blue atom is nitrogen; yellow atom is sulfur; and turquoise atom is sodium. Gray sticks indicate double bonds.
PUBLISHERS RESOURCE GROUP

are organic salts of the carboxylic acid cyclamic acid ($C_6H_{11}NHSO_3H$). The University of Illinois received a patent for the production of cyclamates, which they eventually sold to the DuPont chemical company. DuPont, in turn, sold the patent to Abbott Laboratories, who first marketed the product in 1950. The compound rapidly became very popular as an additive for diet foods and drinks. In most formulations, it is mixed with saccharin, another artificial sweetener. Saccharin is ten times as sweet as sodium cyclamate, but it

leaves a metallic aftertaste that sodium cyclamate helps to mask.

The first evidence about possible health effects of ingesting cyclamates was announced in 1969. Scientists reported that they had fed eighty rats a diet that contained large amounts of Sucaryl®, a product containing saccharin and sodium cyclamate, every day for two years. At the end of that time, twelve of the eighty rats had developed bladder cancer. Based on this study, the U.S. Food and Drug Administration (FDA) decided to ban the use of cyclamates as food additives.

The FDA's decision has long been the subject of considerable debate. Some people praise the agency for banning a substance they believe to be carcinogenic in humans as well as rats. Others feel that the evidence on which the FDA decision was based is not strong enough to require banning of the product. In light of this controversy, a number of additional studies have been conducted on the health effects of the cyclamates. In 1984, the FDA reversed its decision on the compound, having become convinced that no evidence had been produced to support the original 1969 study on which its earlier decision had been made. A year later, the National Academy of Sciences announced the results of its own review of the evidence on the cyclamates. The academy also saw no reason to ban the product as a food additive. Outside the United States, the cyclamates have generally been approved for use as an artificial sweetener.

HOW IT IS MADE

Cyclamic acid is made by the sulfonation of cyclohexylamine. Cyclohexylamine is a six-carbon ring compound with a single amine (-NH₂) group attached to it. Its formula is C₆H₁₁NH₂. Sulfonation is the process by which an -SO₃ group is added to a compound. Sulfonation of cyclohexylamine is accomplished with either sulfur dioxide (SO₂) or sulfamic acid (HOSO₂NH₂).

COMMON USES AND POTENTIAL HAZARDS

Sodium cyclamate and calcium cyclamate are used as artificial sweeteners. Since they have no nutritional value

Interesting Facts

- Sodium cyclamate was nearly the undoing of the sports drink Gatorade. The drink was invented in the late 1960s. The inventors decided to use sodium cyclamate as a sweetener because it had less of a bitter aftertaste than saccharin. One year after Gatorade first appeared, however, the FDA banned

cyclamates as food additives. The inventors quickly reformulated their product, using fructose in place of cyclamates as a sweetener. The drink went on to become an outstanding commercial success, but without the benefit of sodium cyclamate.

(that is, the contain no calories), they can be used in foods and drinks designed for diabetics and dieters. The product also appeals to food processors because it is much sweeter than table sugar. One gram of sodium or calcium cyclamate is as sweet as 30 grams of table sugar, so much less is needed to make a product taste sweet. The cyclamates are also stable to heat, which means that they will not break down if foods in which they are contained are baked or boiled. The compounds have long shelf lives and are inexpensive to make.

Cyclamates are often used in combination with table sugar and other artificial sweeteners. The combination of a cyclamate and saccharin, for example, has the benefit that the two sweeteners cancel out the bitter and metallic aftertastes that each by itself has.

Words to Know

CARCINOGEN A substance that causes cancer in humans or other animals.

FOR FURTHER INFORMATION

“Cyclamate.” Zhonga Hua Fang Da.

<http://www.fangda.com.hk/english/> (accessed on November 8, 2005).

“Low Calorie Sweeteners: Cyclamate.” CalorieControl.org.

<http://www.caloriecontrol.org/cyclamat.html> (accessed on November 8, 2005).

“Sodium Cyclamate.” Hazard Database.

http://www.evol.nw.ru/labs/lab38/spirov/hazard/sodium_cyclamate.html (accessed on November 8, 2005).

See Also L-Aspartyl-L-phenylalanine Methyl Ester;
Saccharin; Sucrose



OTHER NAMES:	Sodium monofluoride
FORMULA:	NaF
ELEMENTS:	Sodium, fluorine
COMPOUND TYPE:	Binary salt (inorganic)
STATE:	Solid
MOLECULAR WEIGHT:	41.99 g/mol
MELTING POINT:	996°C (1824°F)
BOILING POINT:	1704°C (3099°F)
SOLUBILITY:	Moderately soluble in water; insoluble in ethyl alcohol

KEY FACTS

Sodium Fluoride

OVERVIEW

Sodium fluoride (SO-dee-um FLOR-ide) is a colorless to white crystalline solid or powder. It is best known for its role in efforts to prevent tooth decay. It may be added to toothpastes or mouthwashes or to municipal water supplies for this purpose. Although the practice of fluoridating water is now widespread in the United States, it remains the subject of controversy regarding its potential health effects on humans.

HOW IT IS MADE

Sodium fluoride occurs naturally as the mineral villiaumite, although the compound is not produced commercially from that source. Some sodium fluoride is obtained as a byproduct of the manufacture of phosphate fertilizers. In that process, apatite (a form of calcium phosphate that also contains fluorides and/or chlorides) is crushed and treated with sulfuric acid (H_2SO_4). The products of that reaction include phosphoric acid (H_3PO_4), calcium sulfate ($CaSO_4$),

SODIUM FLUORIDE

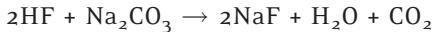
Sodium fluoride. Green atom is fluorine and turquoise atom is sodium. PUBLISHERS RESOURCE GROUP



—



hydrogen fluoride (HF), and silicon tetrafluoride (SiF_4). The hydrogen fluoride and silicon tetrafluoride can then be converted into sodium fluoride. The compound can also be produced by treating hydrogen fluoride with sodium carbonate (Na_2CO_3):



COMMON USES AND POTENTIAL HAZARDS

More than 50 years of research has shown that sodium fluoride and other fluorides are effective in preventing tooth decay. Based on this information, sodium fluoride or some other compound of fluorine is now added to most toothpastes made in the United States. Dentists regularly treat their patients' teeth with fluoride washes to make them more resistant to decay. Most cities and towns in the United States add sodium fluoride or a comparable compound to the municipal water supply to reduce the rate of dental caries (cavities). Some people who live where fluoride is not added to their water supply take sodium fluoride pills to improve their dental health. The American Dental Association and the World Health Organization recommend

Interesting Facts

- Some of the opposition to the fluoridation of water is based on a misunderstanding of the difference between fluorine, the element, and fluorides, compounds of fluorine, such as sodium fluoride and potassium fluoride. Fluorides have very different physical, chemical, and biological

properties from the element fluorine. For example, fluorine gas is a highly toxic gas that reacts violently with most substances, including water. Fluorides, on the other hand, are relatively inert (unreactive) and safe to ingest in small amounts.

fluoridation of drinking water at a level between 0.7 and 1.2 parts per million.

In spite of these trends, opposition to the use of fluorides against tooth decay remains strong in the United States and other parts of the world. Opponents are not convinced that there is sufficient evidence for the claims that fluoridation decreases the rate of tooth decay. They suggest that fluorides may cause cancer and a host of other health problems. And they argue that fluoridating public water supplies removes the choice that individuals should have as to whether or not they want to use fluorides in their dental health program.

Sodium fluoride has a number of commercial and industrial uses in addition to those related to dental health. Those uses include:

- As a wood preservative;
- In the production of certain types of pesticides and as an insecticide for ant and roach control;
- In the preparation of other fluoride salts;
- In electroplating operations;

Words to Know

ELECTROPLATING A process by which a thin layer of one metal is deposited on top of a second metal by passing an electric current through a solution of the first metal.

- In the degassing (removal of gas pockets) during the manufacture of steel;
- For the manufacture of glass and vitreous (glass-like) enamels;
- In detection systems for radiation in the ultraviolet and infrared regions of the electromagnetic spectrum; and
- For disinfecting equipment used in breweries and wineries.

The discussion about fluoridation of public water supplies is sometimes made more difficult because of the fact that sodium fluoride poses some real health hazards to humans. It is a skin, eye, and respiratory irritant that can cause burning if spilled on the skin or in the eyes. If swallowed, it can cause burning of the digestive tract, nausea, vomiting, abdominal pain, stupor, general weakness, tremors, convulsions, collapse, respiratory and cardiac failure, and death. Ingestion of as little as five grams of sodium fluoride can result in death. A condition known as fluorosis is also associated with high doses of sodium fluoride. Fluorosis is characterized by a yellowing and increased brittleness of teeth and bones. Although these conditions are all very serious, they appear only at doses many thousands of times greater than one receives from fluorides in toothpastes and public water supplies.

FOR FURTHER INFORMATION

“Facts about Fluoride.” American Dental Association.
http://www.ada.org/public/topics/fluoride/fluoride_article01.asp (accessed on November 8, 2005).

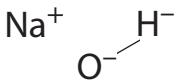
“Fluoride in Drinking Water.” *Science in Dispute*.
Ed. Neil Schlager. Vol. 1. Detroit: Gale, 2002.

“Sodium Fluoride.” Fluoride Action Network.

http://www.flouridealert.com/pesticides/sodium_fluoride_page.htm (accessed on November 8, 2005).

“Sodium Fluoride.” J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/S3722.htm>
(accessed on November 8, 2005).



OTHER NAMES:
Caustic soda; lye;
sodium hydrate;
white caustic

FORMULA:
 NaOH

ELEMENTS:
Sodium, oxygen,
hydrogen

COMPOUND TYPE:
Base (inorganic)

STATE:
Solid

MOLECULAR WEIGHT:
40.00 g/mol

MELTING POINT:
 323°C (613°F)

BOILING POINT:
 1388°C (2530°F)

SOLUBILITY:
Soluble in water,
ethyl alcohol, methyl
alcohol, and glycerol

KEY FACTS

OVERVIEW

Sodium hydroxide (SO-dee-um hye-DROK-side) is a white deliquescent solid commercially available as sticks, pellets, lumps, chips, or flakes. A deliquescent material is one that absorbs moisture from the air. Sodium hydroxide also reacts readily with carbon dioxide in the air to form sodium carbonate. Sodium hydroxide is the most important commercial caustic. A caustic material is a strongly basic or alkaline material that irritates or corrodes living tissue. The compound ranked number 11 among chemicals produced in the United States in 2004.

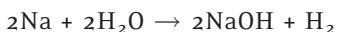
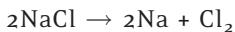
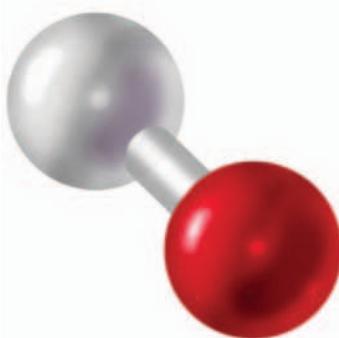
HOW IT IS MADE

Sodium hydroxide is produced commercially simultaneously with chlorine gas by the electrolysis of a sodium chloride solution. In this process, an electric current breaks down sodium chloride into its component elements, sodium and chlorine. The chlorine escapes as a gas, while the sodium metal form reacts with water to form sodium hydroxide:

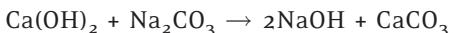
SODIUM HYDROXIDE



Sodium hydroxide. Red atom is oxygen; white atom is hydrogen; and turquoise atom is sodium. PUBLISHERS RESOURCE GROUP



Sodium hydroxide can also be produced easily by means of other chemical reactions. For example, the reaction between slaked lime (calcium hydroxide; $\text{Ca}(\text{OH})_2$) and soda ash (sodium carbonate; Na_2CO_3) produces sodium hydroxide:



None of these alternative methods can compete economically, however, with the preparation by electrolysis.

Interesting Facts

- Solutions of sodium hydroxide are made by adding the solid compound to water, and never water to the solid. The reason is that large amounts of heat are generated when sodium hydroxide dissolves in water. That heat is absorbed by water, but would not be absorbed by solid sodium hydroxide.
- A popular food in Scandinavian countries, lutefisk, is made by soaking dried fish in sodium hydroxide until it turns into a jelly. The jelly is then soaked in water for several days to remove the poisonous lye.

COMMON USES AND POTENTIAL HAZARDS

Sodium hydroxide has a great variety of household and industrial uses. It is the active ingredient in drain cleaners such as Drano® because it breaks up and dissolves the greasy mass that is responsible for drain blockages. It is also an ingredient in many other household products, including oven cleaners, metal polishes, and hair straighteners. Sodium hydroxide is also used in the preparation of homemade and processed foods. It is used in the preparation of soft drinks, chocolate, ice creams, caramel coloring, and cocoa. Hominy, a starchy food similar to grits, is made by soaking corn kernels in a solution of sodium hydroxide in water. Bakers glaze pretzels and German lye rolls with a weak lye solution before baking them. The lye gives baked goods a crisp crust. Some people use lye to cure olives.

The largest single use for sodium hydroxide is in the production of organic compounds from which polymers are made, such as propylene oxide and the ethylene amines, and of the polymers themselves, including the polycarbonates and epoxy resins. About a third of all the sodium hydroxide produced in the United States goes to this application. Another important use of sodium hydroxide is in the pulp and paper industry, where it is used to digest (break down) the raw materials from which pulp and paper are made. About 13 percent of all the sodium hydroxide made in the

Words to Know

ELECTROLYSIS A process in which an electric current is used to bring about chemical changes.

HYDROLYSIS The process by which a compound reacts with water to form two new compounds.

United States goes to this application. Sodium hydroxide is also an important raw material in the manufacture of soap. The method by which soap is made has not changed very much for thousands of years. A fat or oil is added to a boiling solution of sodium hydroxide in water. The fat or oil hydrolyzes into its component parts, glycerol and fatty acids. The sodium hydroxide then reacts with the fatty acids, forming sodium salts. The sodium salt of a fatty acid is a soap. Sodium hydroxide is also an important raw material in the manufacture of inorganic compounds, especially sodium and calcium hypochlorite, sodium cyanide, and a number of sulfur-containing compounds. Some other important uses of sodium hydroxide include:

- In the manufacture of cellophane and rayon;
- As a neutralizing agent during the refining of petroleum;
- In the manufacture of aluminum metal;
- For the refining of vegetable oils;
- As an agent for peeling fruits and vegetables for processing;
- In the extraction of metals from their ores;
- For the processing of textiles;
- In water treatment facilities;
- For etching and electroplating operations; and
- In a wide variety of research laboratory applications.

Sodium hydroxide is one of the most caustic substances known and a strong irritant to the skin, eyes, and respiratory system. Exposure to sodium hydroxide dust, powder, or solid can cause burning of the skin and eyes, with possible permanent damage to one's vision. Ingestion of the compound

causes burning of the mouth, esophagus, and stomach, resulting in nausea, diarrhea, internal bleeding, scarring, and permanent damage to the lungs and gastrointestinal system. More serious results, such as a drop in blood pressure and collapse, are also possible.

FOR FURTHER INFORMATION

“Determination of Acute Reference Exposure Levels for Airborne Toxicants.” [Sodium Hydroxide]. Office of Environmental Health Hazard Assessment, State of California.

http://www.oehha.ca.gov/air/acute_rels/pdf/1310932A.pdf (accessed on November 8, 2005).

“DOW Caustic Soda Solution.” Dow Chemical Company.

<http://www.dow.com/causticsoda/prod/process.htm> (accessed on November 8, 2005).

“Sodium Hydroxide.” International Chemical Safety Cards.

<http://www.cdc.gov/niosh/ipcsneng/nengo36o.html> (accessed on November 8, 2005).

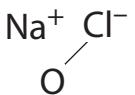
“Sodium Hydroxide.” Medline Plus.

<http://www.nlm.nih.gov/medlineplus/ency/article/002487.htm> (accessed on November 8, 2005).

White, Elaine. “Making Modern Soap with Herbs, Beeswax, and Vegetable Oils.”

<http://www.pioneerthinking.com/soaps.html> (accessed on November 8, 2005).

See Also Chlorine; Potassium Hydroxide

**OTHER NAMES:**

Sodium oxychloride;
hypochlorite; bleach;
chlorine bleach

FORMULA:

NaClO

ELEMENTS:

Sodium, chlorine,
oxygen

COMPOUND TYPE:
Oxy salt (inorganic)**STATE:**

Solid or aqueous
solution; See Over-
view

MOLECULAR WEIGHT:
74.44 g/mol**MELTING POINT:**

Solid NaClO explodes
on heating.

The pentahydrate
(NaClO·5H₂O) is more
stable; its melting
point is 18°C (64°F)

BOILING POINT:

Not applicable;
decomposes

SOLUBILITY:

Soluble in water

K E Y F A C T S**OVERVIEW**

Sodium hypochlorite (SO-dee-um hye-po-KLOR-ite) is the active ingredient in liquid chlorine bleaches, used in the home and many industries to whiten fabric and other materials and to disinfect surfaces and water. The anhydrous compound is very unstable and explodes readily. The pentahydrate is a pale-green crystalline solid that is relatively stable. The compound is usually made available as an aqueous solution that contains anywhere from 3 to 6 percent sodium hypochlorite (for household use) to as high as 30 percent (for industrial applications). In solution form, sodium hypochlorite is quite stable and can be stored for long periods of time out of sunlight.

Sodium hypochlorite decomposes by two mechanisms. In one case, it breaks down to form sodium chloride and sodium chlorate:



In the second case, it breaks down to form sodium chloride and nascent (free single atoms) oxygen:

Sodium hypochlorite. Red atom is oxygen; green atom is chlorine; and turquoise atom is sodium. PUBLISHERS RESOURCE GROUP



Nascent oxygen is a very active form of oxygen that is responsible for the bleaching and disinfectant properties of sodium hypochlorite.

Humans have long made efforts to bleach fabrics. Neither cotton nor linen, two very popular fabrics, are naturally very white, so efforts were made to find ways to convert them to white materials. Those efforts were not very successful until the discovery of chlorine by the Swedish chemist Karl Wilhelm Scheele (1742–1786) in the 1770s. The powerful oxidizing powers of chlorine made it a likely candidate for use as a bleach (although chemists at the time did not understand how bleaching occurred). The first person to take advantage of chlorine's bleaching powers was the French chemist Claude Louis Berthollet (1748–1822), who lived in Javelle, France. Berthollet produced a weak solution of sodium hypochlorite by passing chlorine gas through sodium carbonate. The product had excellent bleaching powers, and it became known by the name of *eau de Javelle* or *eau de Berthollet*. Berthollet's invention came at just the right time. The Industrial Revolution was just getting under way, and the invention of machines like the spinning jenny and the power loom mechanized the commercial production of cotton and linen cloth and rapidly increased the demand for bleaching agents.

HOW IT IS MADE

Sodium hypochlorite is made commercially by passing chlorine gas (Cl_2) through a cold aqueous solution of sodium hydroxide (NaOH):



The pentahydrate then can be extracted by crystallization.

COMMON USES AND POTENTIAL HAZARDS

Sodium hypochlorite is used almost exclusively for one of two purposes: bleaching or purification. The compound is available commercially for household use under a number of trade names, including Antiformin®, B-K Liquid®, Clorox®, Dakin's Solution®, Dazzle®, Hyclorite®, Javelle water, Piochlor®, Purex®, and Saniton Toothbrush Sanitizer®. About two-thirds of the sodium hypochlorite made in the United States is used as a laundry bleach and sanitizer, in restaurants and institutional kitchens for bleaching and sanitation, and for water purification in residential pools and spas. Some examples of the ways in which sodium hypochlorite is used include:

- For the sterilization of milking equipment and containers at dairy farms;
- For the cleaning and sterilization of work surfaces by amateur and professional beer and wine makers;
- As an ingredient in home cleaning agents, such as toilet bowl sanitizers, mold removers, and drain cleaners; and
- As a disinfectant in private water wells to prevent the growth of microorganisms.

About half of all the sodium hypochlorite used for industrial purposes is consumed in municipal and water treatment systems. About a third of the compound production goes to the sterilization of municipal and commercial swimming pools. The remaining sodium hypochlorite is used in commercial, municipal, and industrial cleaning and bleaching operations.

Sodium hypochlorite is a fire and health hazard. It reacts strongly with metals and organic materials. The rate of reaction increases with the concentration of sodium hypochlorite in solution, so industrial and municipal formulations present a greater environmental threat than do household products. One combination of special concern to consumers is the reaction between sodium hypochlorite and compounds that contain ammonia. For example, the combination of household bleach and household ammonia can produce an explosive or flam-

Interesting Facts

- Until the discovery of chlorine, cloth was usually bleached by soaking it in sour milk or buttermilk and letting it sit in the sun. The process often took up to eight weeks and required large “bleaching fields” on which the cloth could be laid out.
- The first attempt to apply chemical principles

to the practice of bleaching was documented in a book on the subject by the Scottish physician Francis Home, published in 1756. Home suggested using a weak solution of sulfuric acid for bleaching, a practice that reduced bleaching time to about 12 hours.

mable mixture. The fumes from this combination can also be harmful, even deadly. Similarly, household bleach should not be used to clean spills that contain urine since urine itself contains ammonia. Sodium hypochlorite is also incompatible with hydrogen peroxide and acidic products.

Sodium hypochlorite is an irritant to the skin, eyes, and respiratory system. It can produce inflammation, burning, and blistering of the skin; burning of the eyes, with subsequent damage to one's vision; and irritation of the gastrointestinal system that can result in stomach pain, nausea, vomiting, coughing, and ulceration of the digestive tract.

Words to Know

ANHYDROUS Lacking water of hydration.

AQUEOUS SOLUTION A solution that consists of some material dissolved in water.

OXIDATION A chemical reaction in which oxygen reacts with some other substance or, alternatively, in which some substance

loses electrons to another substance, the oxidizing agent.

PENTAHYDRATE A form of a crystalline compound that occurs with five molecules of water.

FOR FURTHER INFORMATION

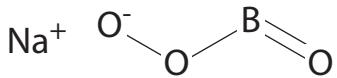
Chalmers, Louis. *Household and Industrial Chemical Specialties*. Vol. 1. New York: Chemical Publishing Co., Inc., 1978.

Fletcher, John, and Don Ciancone. "Why Life's a Bleach (The Sodium Hypochlorite Story)." *Environmental Science & Engineering*. May 1996. Also available online at <http://www.esemag.com/0596/bleach.html> (accessed on November 8, 2005).

"Medical Management Guidelines (MMGs) for Calcium Hypochlorite (CaCl_2O_2) Sodium Hypochlorite (NaOCl)."
Agency for Toxic Substances and Disease Registry.
<http://www.atsdr.cdc.gov/MHMI/mmg184.html> (accessed on January 12, 2006).

"Sodium Hypochlorite." Hill Brothers Chemical Co.
http://hillbrothers.com/msds/pdf/sodium_hypochlorite.pdf
(accessed on November 8, 2005).

"Sodium Hypochlorite." Medline Plus.
<http://www.nlm.nih.gov/medlineplus/ency/article/002488.htm> (accessed on November 8, 2005).



OTHER NAMES:	Perboric acid, sodium salt
FORMULA:	NaBO_3
ELEMENTS:	Sodium, boron, oxygen
COMPOUND TYPE:	Salt (inorganic)
STATE:	Solid
MOLECULAR WEIGHT:	81.80 g/mol
MELTING POINT:	Not available; See Overview
BOILING POINT:	Not applicable
SOLUBILITY:	Soluble in water, with decomposition

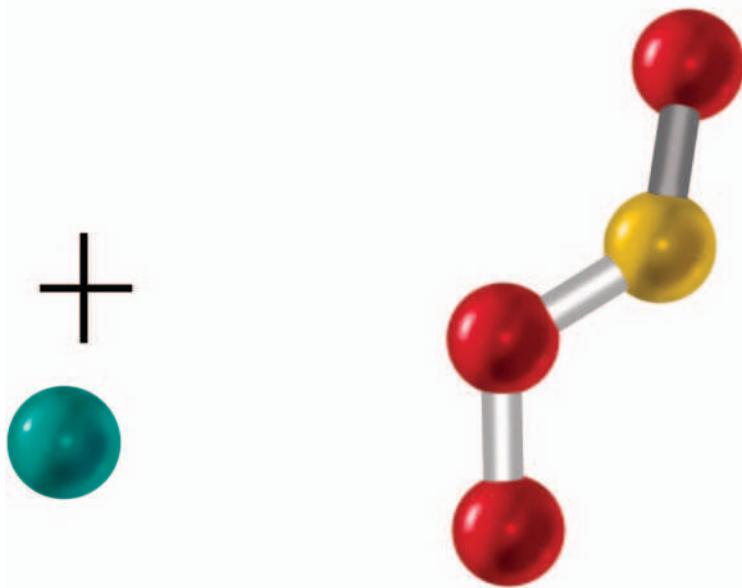
KEY FACTS

Sodium Perborate

OVERVIEW

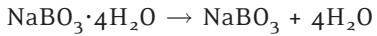
Sodium perborate (SO-dee-um per-BOR-ate) is a white amorphous powder commonly available as the monohydrate ($\text{NaBO}_3 \cdot \text{H}_2\text{O}$) or the tetrahydrate ($\text{NaBO}_3 \cdot 4\text{H}_2\text{O}$). The most frequently available form of the compound, the tetraborate, is a white crystalline solid with a salty taste that melts at 63°C (145°F) and loses its water of hydration when heated above 130°C (270°F). When dissolved in water, all forms of sodium tetraborate decompose to yield hydrogen peroxide (H_2O_2) and sodium borate ($\text{Na}_2\text{B}_4\text{O}_7$). The formation of hydrogen peroxide, which is itself unstable and breaks down to release nascent oxygen (O), makes sodium tetraborate an excellent source of “active” oxygen. The terms *nascent* and *active* refer to individual atoms of oxygen that have a strong tendency to react with other elements and compounds. Since sodium tetraborate is more stable than hydrogen peroxide, it can be used for the same purposes as the peroxide, but is safer and easier to handle. The most common applications of sodium tetraborate are in detergents, bleaches, hair care products, and disinfectants.

Sodium perborate. Red atoms are oxygen; yellow atom is boron; and turquoise atom is sodium. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

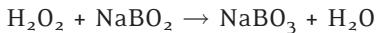


HOW IT IS MADE

Anhydrous sodium perborate can be made by heating the tetrahydrate:



or by reacting hydrogen peroxide with sodium metaborate:



Sodium perborate tetrahydrate is prepared by reacting hydrogen peroxide with borax ($\text{Na}_2\text{B}_4\text{O}_7$).

COMMON USES AND POTENTIAL HAZARDS

Sodium perborates' uses are based on the fact that it is a mild and relatively safe oxidizing agent. An oxidizing agent is a substance that supplies oxygen to other substances. The oxidation of a material may cause bleaching or the destruction of disease-causing microorganisms. For example, sodium perborate is added to some detergents to improve their bleaching capability. It makes the detergents more effective in removing stains, keeping white fabrics white, and preserving the original colors of colored cloth. The compound is

Interesting Facts

- Because of concerns about the hazard that boron compounds may pose to the environment, sodium perborate is being replaced in many applications by a similar but safer oxidizing agent, sodium percarbonate.

also added to some automatic dishwasher powders to improve the product's ability to loosen left-on food and to sterilize dishes, silverware, and cookware. Sodium perborate is also used as an ingredient in a variety of home- and personal-care products, such as deodorants, mouthwashes, denture cleaners, and toothpastes.

Sodium perborate is an irritant to the skin, eyes, and respiratory tract. If ingested, it may cause nausea, vomiting, diarrhea, abdominal pain, and bleeding. Exposure to large quantities of the pure compound can produce severe skin rashes, permanent eye damage, breathing difficulties, unconsciousness, and kidney failure. Despite these concerns, people who use personal- and home-care products containing sodium perborate are at low risk for health problems. The compound degrades during machine washing, and home users rarely ingest, inhale, or come into contact with significant quantities of the compound. The European Union Scientific Committee on Toxicity, Ecotoxicity, and the Environment issued a report in May 2004 stating that manufacturers and consumers do not need to take steps beyond those already in place to protect themselves from exposure to sodium perborate.

FOR FURTHER INFORMATION

"Material Safety Data Sheet." Solvay Interox.

[http://www.solvayinterox.com.au/solvay/uploadfile/
SOLo30%20%20PBST.doc](http://www.solvayinterox.com.au/solvay/uploadfile/SOLo30%20%20PBST.doc) (accessed on November 10, 2005).

"Opinion on the Results of the Risk Assessment of Sodium Perborate." Scientific Committee on Toxicity, Ecotoxicity, and the Environment (CSTEE), 28 May 2004.

[http://europa.eu.int/comm/health/ph_risk/committees/sct/
documents/out225_en.pdf](http://europa.eu.int/comm/health/ph_risk/committees/sct/documents/out225_en.pdf) (accessed on November 10, 2005).

Words to Know

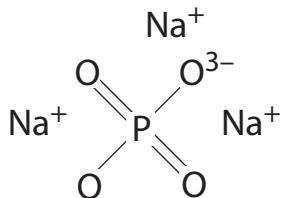
AMORPHOUS Without crystalline structure.

“Sodium Perborate Information.” PAN Pesticides Database.

http://www.pesticideinfo.org/Detail_Poisoning.jsp?Record_Id=PC34416 (accessed on November 10, 2005).

“Sodium Perborate Monohydrate.” Shangyuchem.

http://www.chem-world.com/sodium_perborate.htm (accessed on November 10, 2005).



OTHER NAMES:

See Overview

FORMULA:

Monobasic: NaH_2PO_4 ;

Dibasic: Na_2HPO_4 ;

Tribasic: Na_3PO_4

ELEMENTS:

Sodium, phosphorus,
oxygen

COMPOUND TYPE:

Salt (inorganic)

STATE:

Solid

MOLECULAR WEIGHT:

119.98 to 163.94 g/mol

MELTING POINT:

Not applicable

BOILING POINT:

Not applicable

SOLUBILITY:

Soluble in water;
insoluble in ethyl
alcohol

K
E
Y
F
A
C
T
S

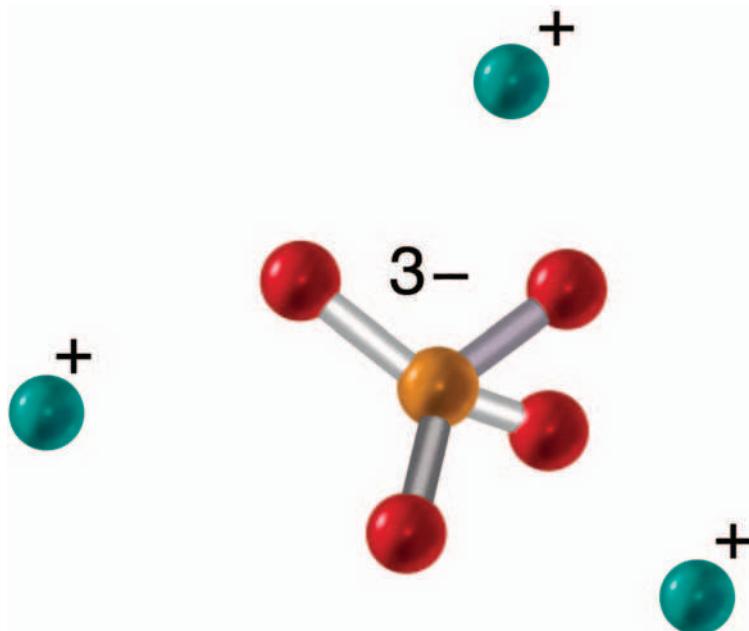
OVERVIEW

The three forms of sodium phosphate are formed when one or more of the three hydrogen atoms in phosphoric acid (H_3PO_4) are replaced by sodium atoms. When one hydrogen is replaced, the monobasic form is produced; replacing two hydrogen atoms results in the formation of the dibasic form; and replacing all three hydrogens results in the formation of tribasic sodium phosphate. All three forms of sodium phosphate are colorless to white crystalline solids or white powders. All may occur as hydrates, such as monobasic sodium phosphate monohydrate and dihydrate ($\text{NaH}_2\text{PO}_4 \cdot \text{H}_2\text{O}$ and $\text{NaH}_2\text{PO}_4 \cdot 2\text{H}_2\text{O}$); dibasic sodium phosphate dihydrate, heptahydrate, and dodecahydrate ($\text{Na}_2\text{HPO}_4 \cdot 2\text{H}_2\text{O}$, $\text{Na}_2\text{HPO}_4 \cdot 7\text{H}_2\text{O}$, and $\text{Na}_2\text{HPO}_4 \cdot 12\text{H}_2\text{O}$); and tribasic sodium phosphate dodecahydrate ($\text{Na}_3\text{PO}_4 \cdot 12\text{H}_2\text{O}$).

The three salts are also known by a number of other names, such as: NaH_2PO_4 : primary sodium phosphate, primary sodium orthophosphate, sodium biprophosphate, MSP; Na_2HPO_4 : secondary sodium phosphate, secondary sodium orthophosphate,

Sodium phosphate. Red atoms are oxygen; orange atom is phosphorus; and turquoise atoms are sodium. Gray sticks indicate double bonds.

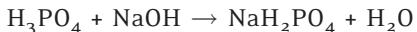
PUBLISHERS RESOURCE GROUP



disodium phosphate, disodium hydrogen phosphate, DSP; Na_3PO_4 ; tertiary sodium phosphate, tertiary sodium orthophosphate, trisodium phosphate, TSP.

HOW IT IS MADE

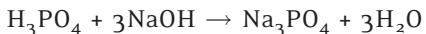
All forms of sodium phosphate are made by treating phosphoric acid with a sodium compound to replace one or more of the hydrogen atoms in the acid. For example, reacting phosphoric acid with sodium hydroxide will make monobasic sodium phosphate:



Reacting phosphoric acid with sodium carbonate will produce the dibasic form of the compound:



And treating phosphoric acid with an excess of sodium hydroxide will result in the formation of the tribasic form of the compound:



Interesting Facts

- Chicken processors often dip whole chickens in a solution of TSP before treating and packaging them. The compound kills bacteria and reduces the risk of food poisoning.

Other methods of preparation are available for all three forms of sodium phosphate.

COMMON USES AND POTENTIAL HAZARDS

The three forms of sodium phosphate have somewhat different uses. Monobasic sodium phosphate is used as a food additive to maintain proper acidity and in baking powders, as a food supplement to provide the phosphorus needed in a person's daily diet, in the treatment of boiler water to reduce the formation of scale on the inner surface of the boiler, and as a feed supplement for cattle and other farm animals.

Dibasic sodium phosphate is used as a food additive to maintain emulsions and proper acidity of food products, in the manufacture of fertilizers, as a food supplement for humans and farm animals, in the treatment of silk, for fire-proofing of wood and paper products, to treat boiler water, in the production of detergents, as a raw material in the manufacture of ceramics, as a mordant for dyeing, and as a cathartic and laxative.

Tribasic sodium phosphate is used in a variety of cleaning agents, such as detergents, industrial cleaning products, and metal cleaners; to treat boiler water; for the tanning of leather; in the manufacture of textiles and paper products; in the purification of sugar; as a dietary supplement for humans and farm animals; in paint removers; and for various photographic purposes.

All three forms of sodium phosphate are mild skin, eye, and respiratory system irritants. Tribasic sodium phosphate has somewhat more serious health consequences than the monobasic or dibasic forms. Exposure to the compounds may produce redness, itching, burns, pain, and blisters on

Words to Know

CATHARTIC A substance that promotes bowel movements.

EMULSION A temporary mixture of two liquids that normally do not dissolve in each other.

MORDANT A substance used in dyeing and printing that reacts chemically with both a dye and the material being dyed to help hold the dye permanently to the material.

the skin; redness, pain, and burning of the eyes; and a burning sensation, coughing, shortness of breath, and a sore throat if ingested. Taken in larger amounts, the tribasic form may cause vomiting, nausea, and diarrhea and may induce shock and collapse. In such cases, immediate medical assistance is required.

FOR FURTHER INFORMATION

“Sodium Phosphate Dibasic Heptahydrate.” Ted Pella, Inc.

http://www.tedpella.com/msds_html/19545msd.htm (accessed on November 10, 2005).

“Sodium Phosphate, Monobasic.” Cornell University.

<http://msds.ehs.cornell.edu/msds/msdsdod/a9/m4481.htm#Section3> (accessed on November 10, 2005).

“Sodium Phosphates.” Agricultural Marketing Service, U.S. Department of Agriculture.

<http://www.ams.usda.gov/nop/NationalList/TAPReviews/sodiumphosphates.pdf> (accessed on November 10, 2005).

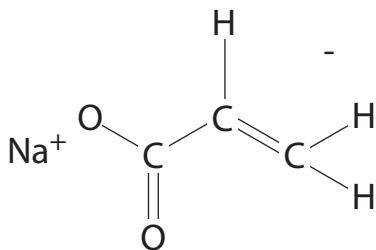
“Trisodium Phosphate.” Household Products Database, National Institutes of Health.

<http://householdproducts.nlm.nih.gov/cgi bin/household/brands?tbl=brands&id=3007033> (accessed on November 10, 2005).

“Trisodium Phosphate (Anhydrous).” International Chemical Safety Cards.

<http://www.inchem.org/documents/icsc/icsc/eics1178.htm> (accessed on November 10, 2005).

“USDA Approves Phosphate to Reduce Salmonella in Chicken.” *Environmental Nutrition* (February 1993): 3.



OTHER NAMES:

Acrylic acid polymer
sodium salt; poly-
acrylic acid sodium
salt; PAAS

FORMULA:
 $-[-\text{CH}_2-\text{CH}(\text{COONa})]_n-$

ELEMENTS:
Carbon, hydrogen,
oxygen, sodium

COMPOUND TYPE:
Polymer of an
organic salt

STATE:
Solid

MOLECULAR WEIGHT:
Varies

MELTING POINT:
Not applicable

BOILING POINT:
Not applicable

SOLUBILITY:
Insoluble in water;
swells in water

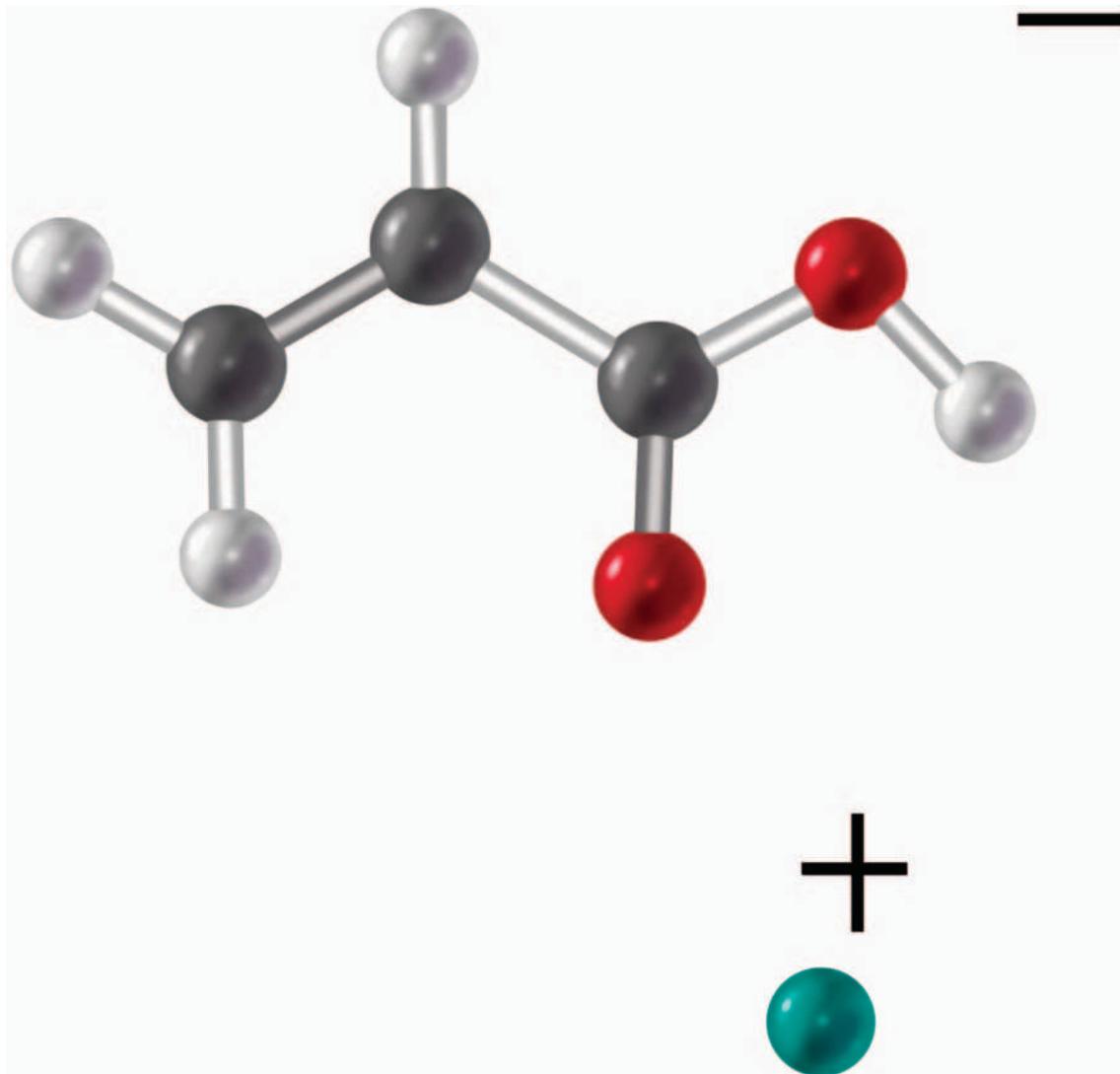
KEY FACTS

OVERVIEW

Sodium polyacrylate (SO-dee-um pol-ee-AK-ruh-late) is an odorless, grainy white powder. Its most impressive property is its ability to absorb large amounts of fluid, up to 800 times its volume of distilled water and lesser amounts of other liquid mixtures. This property accounts for one of its primary applications, in the manufacture of disposable diapers. Diapers made from sodium polyacrylate are able to absorb up to 30 grams of urine for each gram of diaper.

HOW IT IS MADE

Sodium polyacrylate is produced by the reaction between acrylic acid ($\text{H}_2\text{C}=\text{CHCOOH}$) and its sodium salt ($\text{H}_2\text{C}=\text{CHCOONa}$). The product of this reaction is a long-chain copolymer consisting of alternate units of acrylic acid and sodium acrylate. A copolymer is a polymer made of two different monomers, in this case, acrylic acid and sodium acrylate. What makes this polymer different from most other



Sodium polyacrylate. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; turquoise atom is sodium. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

kinds of polymers is that adjacent polymer chains are able to cross link with each other. The hydrogen on a carboxyl group ($-\text{COOH}$) on one chain reacts with a double bond ($-\text{C}=\text{C}-$) on an adjacent chain, forming a link that holds the two chains together. Cross-linking occurs at many points in the polymer, resulting in the formation of a mesh-like web consisting of polymer chains.

When water is added to the polymer, it forces carboxyl groups away from each other, forcing the mesh to open up

Interesting Facts

- Sodium polyacrylate was first developed by researchers for the National Aeronautics and Space Administration. The material was used for diapers worn by astronauts while they were on long space trips.

and make space for water molecules to fill the gaps in the polymer. As more water is added, the carboxyl groups stretch even farther apart, making room in turn for yet more water molecules to be absorbed by the polymer. If the polymer is allowed to dry out, water molecules leave gaps in the compound, the empty spaces between carboxyl groups collapse, and the polymer returns to its original size. It can then be stored and re-used any number of times.

COMMON USES AND POTENTIAL HAZARDS

The primary use of sodium polyacrylate is in the manufacture of baby diapers. The need for some sort of disposable diaper first arose during World War II because of a shortage of cotton, from which cloth diapers are made. The compound now has a number of other applications. For example, it is used to pack poultry, red meat, fish, fresh-cut fruits and vegetables, and fresh whole berries to keep them moist and fresh. Fluids from washing these foods during processing may accumulate inside a package and provide an environment for the growth of bacteria that cause foods to spoil. A packaging material made of cellulose and sodium polyacrylate absorbs these fluids and prevents them from being squeezed out of the package.

Sodium polyacrylate is also used as a thickening agent in medical gels used to treat bed sores, which are open wounds that develop when a person is bed-ridden for too long. The compound is also added to detergents and to potting soils to help retain water. The compound is now being used in some parts of the world where there is insufficient rain to allow

Words to Know

COPOLYMER A polymer that consists of two or more different monomers.

POLYMER A compound consisting of very large molecules made of one

or two small repeated units called monomers.

crops or lawns to grow. It absorbs moisture when rain does fall and holds it in place until plants can absorb the water.

In industry, sodium polyacrylate is used in filtration units that remove water from airplane and automotive fuel. It is also used as a thickening agent in coatings and adhesives used in the upholstery, drapery, carpet, paper, paint, wallpaper, printing, and textile industries. The compound is also used to thicken certain liquid products applied by spraying, such as cleaning products. Finally, it is sometimes used to prevent fluid loss in oil wells.

Sodium polyacrylate is a mild irritant to the skin, eyes, and respiratory tract. It may cause redness, itching, and pain on the skin or in the eyes; and coughing, shortness of breath, and inflammation of the respiratory tract. Questions have been raised about possible health effects on babies who wear disposable diapers containing sodium polyacrylate. Some people suggest that a baby's tender skin may be more sensitive to the irritation caused by sodium polyacrylate than the skin of an adult. The compound was removed from tampons in 1985 because some women who left their tampons in place too long experienced unacceptable levels of irritation caused by sodium polyacrylate in the product.

FOR FURTHER INFORMATION

Allison, Cathy. "Disposable Diapers: Potential Health Hazards?" BCParent Online.

http://www.bcparent.com/articles/baby_talk/disposable_diapers.html (accessed on November 10, 2005).

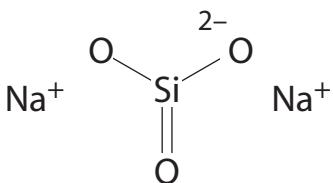
"Environmental Assessment." U.S. Food and Drug Administration, Center for Food Safety and Applied Nutrition.

<http://www.cfsan.fda.gov/~acrobat2/fneao427.pdf> (accessed on November 10, 2005).

Mebane, Robert C., and Thomas R. Rybolt. *Plastics and Polymers*.
New York: Twenty First Century, 1995.

“Sodium Polyacrylate.” Center for Advanced Microstructures and
Devices, Louisiana State University.
http://www.camd.lsu.edu/msds/s/sodium_polyacrylate.htm
(accessed on November 10, 2005).

“Sodium Polyacrylate.” Flinn Scientific Company/Department of
Chemistry, Iowa State University.
http://avogadro.chem.iastate.edu/MSDS/Na_polyacrylate.pdf
(accessed on November 10, 2005).



OTHER NAMES:

Sodium metasilicate; soluble glass; water glass

FORMULA:

Varies; often represented as Na_2SiO_3 ; see Overview

ELEMENTS:

Sodium, silicon, oxygen

COMPOUND TYPE:

Salt (inorganic)

STATE:

Ranges from solid to liquid

MOLECULAR WEIGHT:

Varies; sodium metasilicate: 122.06 g/mol

MELTING POINT:

Varies; sodium metasilicate: 1089°C (1992°F)

BOILING POINT:

Not applicable

SOLUBILITY:

Soluble in cold water; reacts chemically with hot water

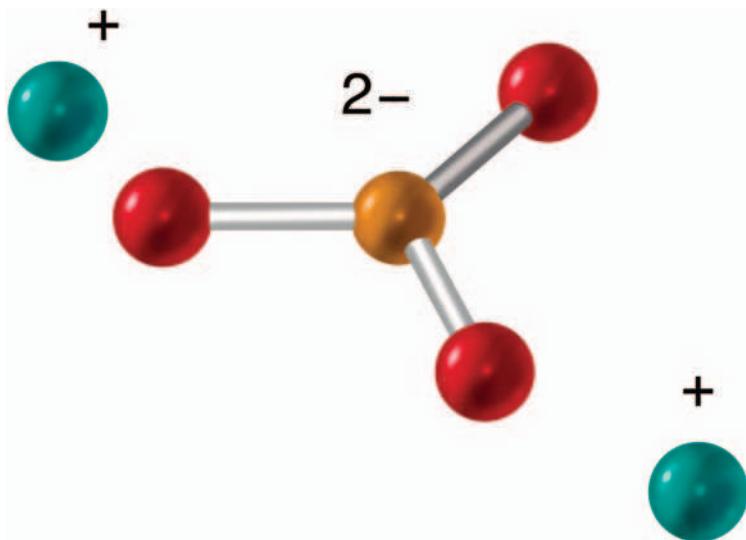
**K
E
Y
F
A
C
T
S**

Sodium Silicate

OVERVIEW

Sodium silicate (SO-dee-um SILL-uh-kate) is a generic term that refers to a series of compounds with varying proportions of sodium oxide (Na_2O) and silicon dioxide (SiO_2). The term is sometimes used synonymously with a specific form of the compound called sodium metasilicate, with the chemical formula Na_2SiO_3 . Other forms typically range from $\text{Na}_2\text{O} \cdot 3.75\text{SiO}_2$ to $2\text{Na}_2\text{O} \cdot \text{SiO}_2$ with one or more molecules of water of hydration. Sodium silicate is a white amorphous (without crystalline structure) solid with the physical properties listed above. Other forms of sodium silicate range from white powders with varying degrees of solubility in water, to greenish, glass-like solids, to liquids with varying degrees of viscosity. Viscosities vary over a range of nearly a million, from about 0.5 poise to more than 600,000 poise. Poise is the unit of measurement for viscosities. Other physical properties of the sodium silicates also vary considerably depending on the relative proportions of sodium oxide and silicon dioxide present. The sodium

Sodium silicate. Red atoms are oxygen; orange atom is silicon; and turquoise atoms are sodium. Gray stick indicates double bond. PUBLISHERS RESOURCE GROUP



silicates are sometimes referred to as the simplest form of glass.

HOW IT IS MADE

Sodium silicates are made by fusing (melting) sand (silicon dioxide) and soda ash (sodium carbonate) or sodium hydroxide in a gas-fired open hearth furnace, somewhat similar to the furnaces used in the manufacture of steel. The products of this reaction are lumps of sodium silicate that are broken apart and dissolved in a stream of hot steam. The proportions of sand and soda ash used, the temperature of the reaction, and the amount of water that remains in the final product all determine the physical properties of the final product.

COMMON USES AND POTENTIAL HAZARDS

About 1.1 million metric tons (1.3 million short tons) of sodium silicate was produced in the United States in 2004. The primary application for the compound is in the manufacture of soaps and detergents. It improves the cleaning ability of these products and is less damaging to metal components of dishwashers and washing machines than other ingredients of soaps and detergents. Sodium silicate

Interesting Facts

- One of the early uses of sodium silicate was as a preservative for eggs. Eggs were soaked in solutions of sodium silicate, or the compound was painted on the egg shells. The sodium silicate filled the pore in the egg shell, preventing bacteria from

entering the egg and causing it to spoil. The process was very effective and preserved eggs for up to nine months. More efficient and less expensive methods of preserving eggs are now available.

is also used as a water softener, used by itself or as an ingredient in detergents.

The next most important applications of sodium silicate are as catalysts and in the pulp and paper industry. Catalysts are materials that increase the rate of a chemical reaction without undergoing any change in their own chemical structure. In the pulp and paper industry, sodium silicate is used to bleach raw pulp and help remove ink from scrap paper being reprocessed. The compound is also used in sealants and adhesives. Some other applications of sodium silicate include:

- For the purification of water in municipal and industrial water treatment plants;
- For the fireproofing of fabrics;
- As an anti-caking agent in food products;
- As an additive in cements, where it helps the cement set more quickly;
- In the manufacture of other compounds of silicon by the chemical industry;
- In fluids used to lubricate drilling instruments;
- As a liner for chemical and industrial equipment, such as furnaces used to make steel;

Words to Know

MUCOUS MEMBRANES The soft tissues that line the breathing and digestive passages.

VISCOUS Describing a syrupy liquid that flows slowly.

- In the processing of ores from which metals are extracted; and
- As a binder on grindstones and abrasive wheels.

Sodium silicates are strong irritants to the skin, eyes, and respiratory system. Prolonged exposure to sodium silicate dust, powder, or liquid may cause inflammation of the skin, eyes, nose, and throat. More serious symptoms may include difficulty in swallowing, burns inside the stomach, damage to the mucous membranes, rapid heartbeat, hypertension, shock, severe damage to the lining of the gastrointestinal tract, various types of cancer, and death. These hazards are of concern primarily to workers who come into contact with sodium silicate in solid or liquid form in the workplace.

FOR FURTHER INFORMATION

Higgins, Kevin T. "Simplified Food Oil Refining." *Food Engineering* (February 1, 2003). Also available online at http://www.foodengineeringmag.com/CDA/ArticleInformation/features/BNP_Features_Item/0,6330,94941,00.html (accessed on November 10, 2005).

"Practical Uses for Sodium Silicate." The Chemistry Store.com. http://www.chemistrystore.com/sodium_silicate_uses.htm (accessed on November 10, 2005).

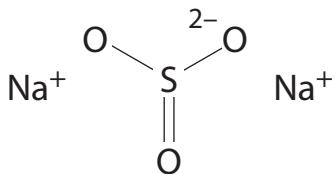
"Silicate Chemistry." PQ Corporation. http://www.pqcorp.com/technicalservice/understanding_silicatesolchem.asp (accessed on November 10, 2005).

"Sodium Metasilicate." International Programme on Chemical Safety. <http://www.inchem.org/documents/pims/chemical/>

pim500.htm#SectionTitle:1.3%20%20Synonyms (accessed on November 10, 2005).

“Sodium Silicates.” Chemical Land 21.

<http://www.chemicalland21.com/arokorhi/industrialchem/inorganic/SODIUM%20SILICATE.htm> (accessed on November 10, 2005).



	OTHER NAMES:	Disodium sulfite
	FORMULA:	Na_2SO_3
	ELEMENTS:	Sodium, sulfur, oxygen
	COMPOUND TYPE:	Salt (inorganic)
	STATE:	Solid
	MOLECULAR WEIGHT:	126.04 g/mol
	MELTING POINT:	Not applicable; decomposes
	BOILING POINT:	Not applicable
	SOLUBILITY:	Soluble in water; slightly soluble in ethyl alcohol

KEY FACTS

Sodium Sulfite

OVERVIEW

Sodium sulfite (SO-dee-um SUL-fite) is a white powder or crystalline solid with no odor but a slightly salty taste. The compound is stable in dry air, but tends to decompose in moist air to produce sulfur dioxide (SO_2) and sodium hydroxide (NaOH). The compound has a variety of uses as a food preservative and in the paper and pulp industry.

HOW IT IS MADE

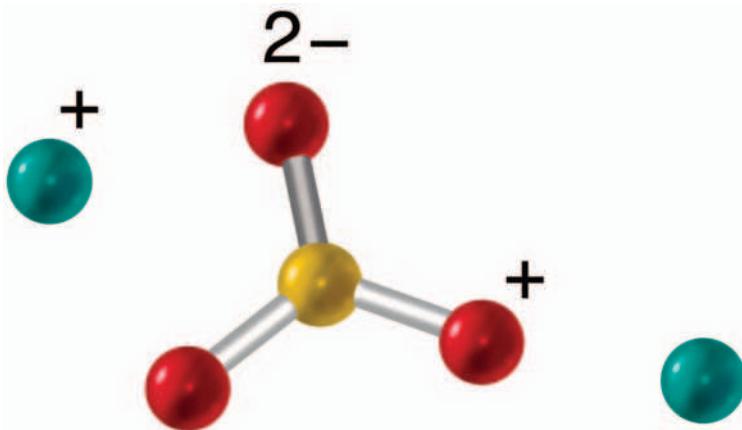
Sodium sulfite can be prepared by reacting sulfur dioxide, soda ash (sodium carbonate; Na_2CO_3), and water. The product of this reaction is sodium bisulfite (NaHSO_3), which is then treated with excess soda ash to obtain sodium sulfite. The compound can also be obtained as a byproduct in the preparation of phenol ($\text{C}_6\text{H}_5\text{OH}$).

COMMON USES AND POTENTIAL HAZARDS

Sodium sulfite is an essential chemical in the pulp and paper industry. Just over half of all the sodium sulfite made

SODIUM SULFITE

Sodium sulfite. Red atoms are oxygen; yellow atom is sulfur; and turquoise atoms are sodium. PUBLISHERS RESOURCE GROUP



in the United States is used by the pulp and paper industry. The compound acts as a pulping agent for wood, rags, and straw. A pulping agent is a substance that breaks down raw materials and converts them into the pulp from which paper is made. Sodium sulfite is also used to remove excess chlorine used to bleach wood pulp and other raw materials needed in the production of paper.

The second largest application of sodium sulfite is in water and wastewater treatment plants, where it is used to react with and neutralize excess chlorine used in the water and wastewater treatment processes. The third most important application of sodium sulfite is in photography. The compound is used in the developing process, and it acts as a preservative for the final picture produced. Sodium sulfite is still used as a food preservative also, although the conditions under which it can be added are somewhat limited. In addition to its hazards among individual with allergies, sodium sulfite destroys both vitamins B₁ and E, meaning that it cannot be added to foods that contain these vitamins. It is still widely used, however, in the wine-making industry for the control of bacteria involved in the wine-making process.

As noted above, a significant number of people are allergic to sodium sulfite. In addition to this health hazard, the compound can be an irritant to the skin, eyes, and respiratory tract. It can cause inflammation of the skin and eyes, irritation of the nose and throat, problems with breathing, and stomach upset. With the level at which most people come

Interesting Facts

- Sodium sulfite has been used as a food preservative for many years. However, in 1986 the U.S. Food and Drug Administration (FDA) banned the use of sodium sulfites for certain types of food. The agency had discovered that about one in a hundred people are sensitive to sodium sulfite. Thirteen deaths and more than 500 allergic

reactions resulting from exposure to sodium sulfite had been reported to the FDA. The agency now prohibits the use of sodium sulfite as a preservative on raw fruits and vegetables. Processed foods that contain sodium sulfite must include a notice to that effect on the food label.

into contact with the compound, however, it poses little threat to a person's health.

FOR FURTHER INFORMATION

"Sodium Sulfite." Esseco General Chemical.

<http://www.genchemcorp.com/pdf/msds/Sodium%20Sulfite,%20EGC%20%204%20.pdf> (accessed on November 12, 2005).

"Sodium Sulfite." J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/s5066.htm> (accessed on November 12, 2005).

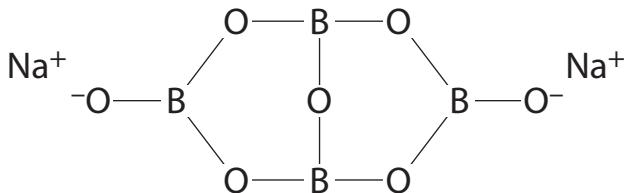
"Sodium Sulfite." Solvay Chemicals.

http://www.solvaychemicals.us/pdf/Sodium_Sulfite/SODSULF.pdf (accessed on November 12, 2005).

"Sodium Sulfite Photographic Grade." Center for Advanced Micro structures and Devices, Louisiana State University.

http://www.camd.lsu.edu/msds/s/sodium_sulfite.htm (accessed on November 12, 2005).

See Also Sulfur Dioxide



OTHER NAMES:

Sodium borate;
sodium pyroborate;
disodium tetraborate;
borax

FORMULA:

$\text{Na}_2\text{B}_4\text{O}_7$ or
 $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$; see
 Overview

ELEMENTS:

Sodium, boron,
oxygen

COMPOUND TYPE:
Salt (inorganic)

STATE:
Solid

MOLECULAR WEIGHT:

$\text{Na}_2\text{B}_4\text{O}_7$: 201.22 g/
 mol; $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$:
 381.37 g/mol

MELTING POINT:
 $\text{Na}_2\text{B}_4\text{O}_7$: 743°C
 (1370°F);

$\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$:
 decomposes at about
 75°C (170°F)

BOILING POINT:
 $\text{Na}_2\text{B}_4\text{O}_7$: 1575°C
 (2867°F);

$\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$: not
 applicable

SOLUBILITY:
Soluble in water

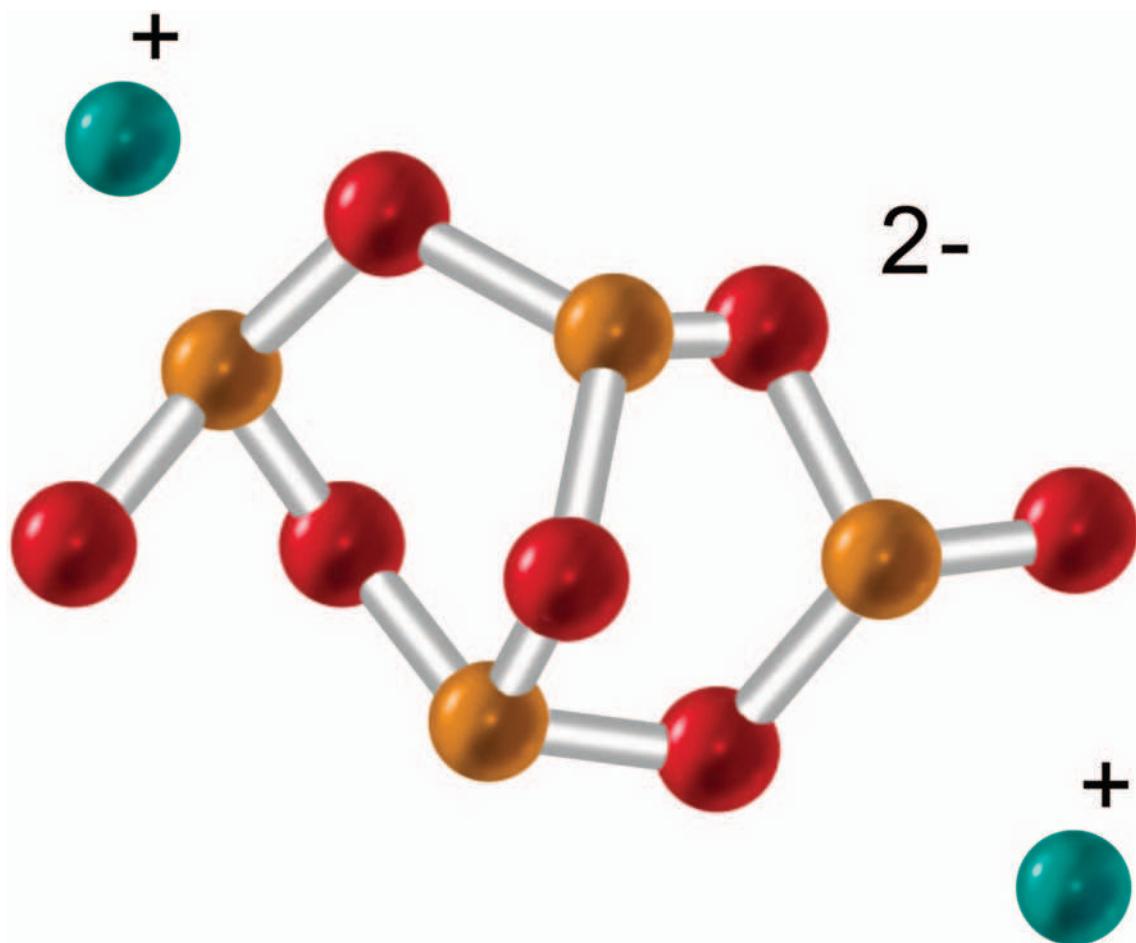
K E Y F A C T S

OVERVIEW

Sodium tetraborate (SO-dee-um tet-ruh-BOR-ate) is a term used for either the anhydrous or hydrated form of the compound with the formula $\text{Na}_2\text{B}_4\text{O}_7$. The decahydrate ($\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$) is also referred to as borax. Borax also occurs without water of hydration and in that form is known as anhydrous borax.

Some historians think that borax may have been known as long as 4,000 years ago. Since ancient people did not use the term, however, there is considerable doubt as to the authenticity of these claims. The compound was certainly in use as far back as about 800 BCE when the compound was being used in the Chinese and Islamic civilization for making glass and in jewelry work. The substance was very expensive, however, and it was not widely used in Europe until the Middle Ages. Borax became more commonly used after extensive deposits of its naturally occurring form were found in the United States. The first of those deposits was discovered in Nevada in 1879, although the largest deposits were later found in the desert regions of southern California. Today, the

SODIUM TETRABORATE



Sodium tetraborate decahydrate. Red atoms are oxygen; white atoms are hydrogen; orange atoms are boron; and turquoise atoms are sodium. PUBLISHERS RESOURCE GROUP

largest reserves of borax are found near the town of Boron, California, and at Borax Lake, California. The compound is also mined in Tibet, Russia, Chile, and Turkey.

Sodium tetraborate is an odorless white crystalline solid or powder. The hydrated form loses its water of hydration when heated and then fuses (melts) to form a glass-like solid at higher temperatures.

HOW IT IS MADE

Sodium tetraborate occurs naturally as the minerals tincal (pronounced “tinkle;” $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$) and kernite

Interesting Facts

- Perhaps the best known commercial form of sodium tetraborate is called Twenty-Mule-Team® Borax. The name comes from the fact that the first borax mines in California were located 165 miles from the nearest train station in Mojave, California. The mined borax was transported that distance in wagons pulled by 20 mules. Each wagon cost \$900 to build, weighed 14,500 kilograms (32,000 pounds), and had wheels 2 meters (7 feet) in diameter. The trip took about twenty days, often in temperatures as high as 45°C (113°F). Over the six years during which mule teams were used, about nine million kilograms (20 million pounds) of borax were moved from mine to railway station. In 1896 the Pacific Coast Borax Company (later, U.S. Borax) took the twenty-mule teams as their corporate symbol.

$(Na_2B_4O_7 \cdot 4H_2O)$. Ores containing these minerals are crushed, washed, and processed to obtain the decahydrate of high purity. Anhydrous sodium tetraborate can be obtained by heating the decahydrate. Sodium tetraborate can also be obtained by processing other minerals that contain borates, such as ulexite ($NaCaB_5O_9 \cdot 8H_2O$) and colemanite ($Ca_2B_6O_{11} \cdot 5H_2O$).

COMMON USES AND POTENTIAL HAZARDS

The primary use of sodium tetraborate is in the manufacture of glass products. About 43 percent of all the compound used in the world goes to this application. Glass made with sodium tetraborate is very strong and heat resistant. The well-known Pyrex® brand of glass is made with sodium tetraborate. Today, the largest single use of borax glass is in the manufacture of fiberglass insulation and fiberglass textiles.

The next most important use of sodium tetraborate is in the manufacture of soaps, detergents, and personal care products. Some well known products that contain borax

Words to Know

ANHYDROUS Describing a compound that lacks any water of hydration.

DECAHYDRATE Form of a compound that exists with ten molecules of water.

FLUX A material that aids the processes of welding and soldering (joining) metals.

WATER OF HYDRATION Water that has combined with a compound by some physical means.

include 20-Mule-Team® Borax all-purpose cleaner, 20-Mule-Team® Borax Laundry Booster, Borateem® stain remover, and Boraxo® powdered hand soap. Some other uses of sodium tetraborate include:

- As a flame-retardant and fungicide for wood products;
- In the production of enamel, porcelain, glazes, enamels, and frits (specialized types of glass);
- In the manufacture of fertilizers and herbicides;
- As additives for certain kinds of polymers;
- As a flux for smelting and soldering metals;
- In the preparation of rust inhibitors; and
- In certain photographic processes.

Sodium tetraborate is a mild irritant to the skin, eyes, and respiratory tract. It can cause inflammation of the skin, eyes, nose, throat, and lungs. If ingested, it can cause nausea, vomiting, diarrhea, and abdominal pain. There is some evidence that ingestion of sodium tetraborate can cause reproductive problems in laboratory animals, although similar effects have not been seen in humans. Swallowing large quantities of sodium tetraborate can have serious health consequences, especially for young children. The fatal dose of sodium tetraborate for young children is about five grams (0.2 ounce) of the compound.

FOR FURTHER INFORMATION

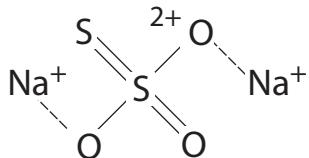
"Borax (Hydrated Sodium Borate)." Amethyst Galleries, Inc.
<http://mineral.galleries.com/minerals/carbonat/borax/borax.htm> (accessed on November 12, 2005).

“Rio Tinto Borax.” Borax.

<http://www.borax.com/index.html> (accessed on November 12, 2005).

“Sodium Tetraborate.” International Programme for Chemical Safety.

<http://www.inchem.org/documents/icsc/icsc/eics1229.htm> (accessed on November 12, 2005).



OTHER NAMES:
Sodium hyposulfite;
thiosulfic acid
sodium salt

FORMULA:
 $\text{Na}_2\text{S}_2\text{O}_3$

ELEMENTS:
Sodium, sulfur, oxygen

COMPOUND TYPE:
Salt (inorganic)

STATE:
Solid

MOLECULAR WEIGHT:
158.11 g/mol

MELTING POINT:
Decomposes at about
100°C (200°F)

BOILING POINT:
Not applicable

SOLUBILITY:
Soluble in water;
insoluble in ethyl
alcohol

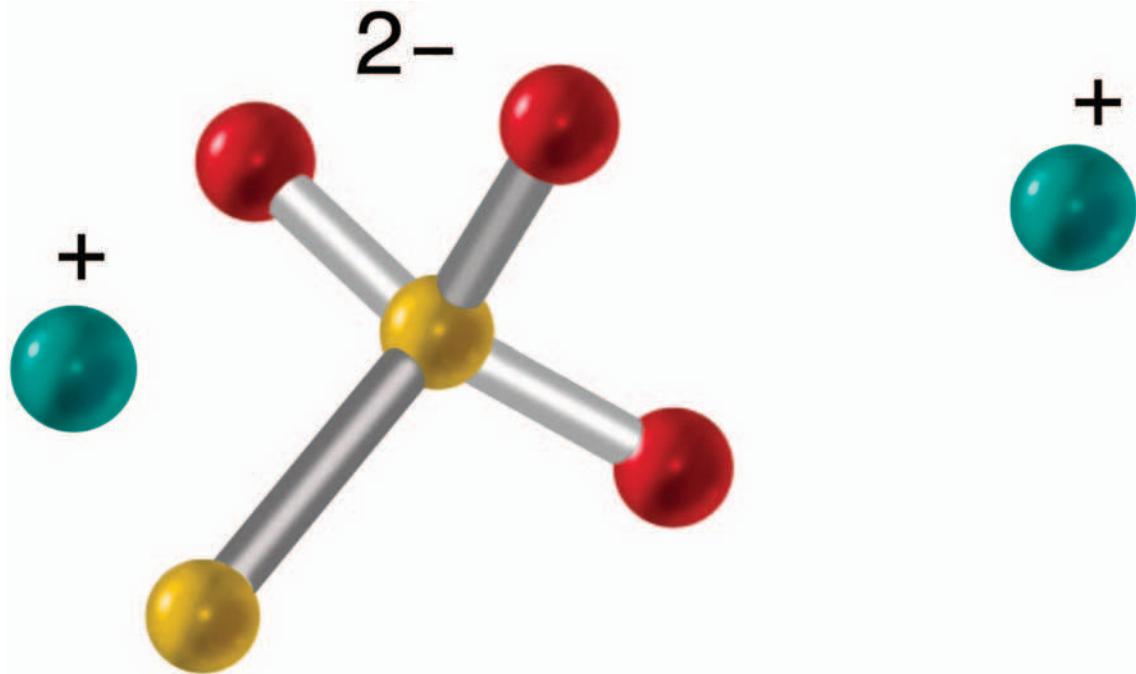
KEY FACTS

OVERVIEW

Sodium thiosulfate (SO-dee-um THYE-oh-SUL-fate) is a colorless to white crystalline solid or powder with no odor and a cooling, bitter taste. The compound usually occurs in the form of the pentahydrate ($\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$). Sodium thiosulfate is an antichlor, a compound that reacts with and neutralizes excess chlorine used in some industrial, commercial, or other applications. Two other popular antichlors are sodium bisulfite (NaHSO_3) and sodium sulfite (Na_2SO_3). Some of the most important applications of antichlors are in the production of pulp and paper and in the textile industry. After pulp, paper, or a textile has been treated with chlorine to bleach the material, an antichlor such as sodium thiosulfate is added to remove any remaining chlorine from the reaction vat. The antichlor itself is then removed by washing with water.

HOW IT IS MADE

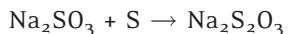
Large amounts of sodium thiosulfate are obtained as the byproducts of other industrial reactions, such as the production



Sodium thiosulfate. Red atoms are oxygen; yellow atoms are sulfur; and turquoise atoms are sodium. Gray sticks indicate double bonds.

PUBLISHERS RESOURCE GROUP

of sodium sulfide (Na_2S) and dyes that contain sulfur. The compound can also be produced directly by adding powdered sulfur to a solution of sodium sulfite and heating the reactants:



COMMON USES AND POTENTIAL HAZARDS

In addition to its use as an antichlor, sodium thiosulfate can, itself, also be used as a bleach for paper, pulp, bone, straw, ivory, and other materials. Its other major application is in photography, where it is used as a fixing agent. A fixing agent is a chemical that reacts with silver bromide and silver chloride on a photographic film that has not been exposed. The silver bromide and silver chloride are then washed away, leaving behind and “fixed in place” the silver that has been produced by exposure to light. Smaller amounts of sodium thiosulfate are used for other purposes, such as:

- A food additive for the purpose of maintaining the proper acidity of a food product or sequestering

Interesting Facts

- An important, but seldom used, application of sodium thiosulfate is as an antidote for cyanide poisoning.

(capturing and holding) unwanted materials in the food;

- For certain medical purposes, such as the treatment of fungal infections of the skin in humans and, specifically, for the treatment of ringworm in animals;
- In hide tanning and dyeing procedures that use compounds of the element chromium;
- For the extraction of silver metal from its ores; and
- In the analysis of the composition of chemical mixtures.

Sodium thiosulfate is an irritant to the skin, eyes, and respiratory tract, although such effects tend to be mild unless one is exposed to large quantities of sodium thiosulfate dust, mist, or solutions. The compound can also be toxic if ingested, causing irritation of the gastrointestinal tract. The amount of sodium thiosulfate permitted in foods is 0.1 percent. No serious long-term effects of exposure to the compound are known.

Words to Know

ANTICHLOR A chemical that reacts with excess chlorine used for purification, disinfecting, or some other purpose.

FOR FURTHER INFORMATION

“Sodium Thiosulfate.” Esseco General Chemistry.

<http://www.genchemcorp.com/pdf/msds/Sodium%20Thiosulfate,%20EGC%20%20403.pdf> (accessed on November 12, 2005).

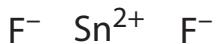
“Sodium Thiosulfate, Pentahydrate.” Boehringer Ingelheim.

http://www.boehringer-ingelheim.ca/vetmedica/msds_sodium.asp (accessed on November 12, 2005).

“Sodium Thiosulfate (Systemic).” Drugs.com.

http://www.drugs.com/cons/Sodium_Thiosulfate.html (accessed on November 12, 2005).

See Also Sodium Sulfite



OTHER NAMES:
Tin(II) fluoride; tin difluoride
FORMULA:
SnF_2
ELEMENTS:
Tin, fluorine
COMPOUND TYPE:
Binary salt (inorganic)
STATE:
Solid
MOLECULAR WEIGHT:
156.71 g/mol
MELTING POINT:
213°C (415°F)
BOILING POINT:
850°C (1560°F)
SOLUBILITY:
Soluble in water; insoluble in ethyl alcohol, ether, and chloroform

K E Y F A C T S

OVERVIEW

Stannous fluoride (STAN-us FLOR-ide) is a lustrous, white crystalline solid with a salty and bitter taste. The only application of any consequence for stannous fluoride is as an additive in toothpastes. Its function is to strengthen tooth enamel and reduce the rate of tooth decay.

The first hints that fluorides might be helpful in preventing tooth decay were reported as early as 1901. Fluorides are binary compounds of the element fluorine, compounds containing one element other than fluorine. Sodium fluoride (NaF), potassium fluoride (KF), and stannous fluoride are typical fluorides. But solid scientific evidence for the role of fluorides in preventing tooth decay was not available until the 1940s. Even then, the idea of adding fluorides to toothpastes to reduce the risk of tooth decay was hampered by researchers' inability to find suitable fluoride compounds to use for this purpose. One research team working on this problem was that of Joseph C. Muhler (1923–1996), then an undergraduate student at Indiana University, and one of his

2+

Stannous fluoride. Green atoms are tin and turquoise atom is fluorine. PUBLISHERS RESOURCE GROUP

instructors, Harry G. Day (1906–). Muhler and Day discovered that stannous fluoride was significantly more effective at preventing cavities than other fluorides they had tested. Furthermore, the compound could be incorporated into toothpastes without affecting their other properties. Their first research paper on the subject was published in 1950. The information discovered by Muhler and Day was sold to the Proctor & Gamble company, which released the first toothpaste containing stannous fluoride—Crest®—in 1955.

HOW IT IS MADE

Stannous fluoride is made commercially by reacting stannous oxide (SnO) with hydrofluoric acid (H_2F_2) in an oxygen-free environment.

COMMON USES AND POTENTIAL HAZARDS

The primary use of stannous fluoride is in toothpastes, dental rinses, and fluoride treatments for teeth. Fluorides prevent tooth decay in two ways. First, fluorides kill bacteria that cause tooth decay. Second, fluorides react with other chemicals in the mouth to make new enamel to replace enamel destroyed by bacteria or worn away by mechanical processes. Studies suggest that stannous fluoride also has other beneficial effects, including preventing gingivitis (inflammation of the gums) and bad breath.

Stannous fluoride does have some disadvantages as an additive in tooth care products. For example, it has a somewhat bitter and unpleasant taste. Today, a number of

Interesting Facts

- The effects of fluorides in preventing tooth decay were first observed by a young dentist named Frederick S. McKay in Colorado Springs, Colorado, in 1901. McKay noticed that many of his patients had unattractive brown stains on their teeth. But the rate

of tooth decay among these patients was very low. He concluded that both the brown stains and the low rate of tooth decay were caused by unusually high levels of fluorides in the local water supply.

alternatives have been developed for stannous fluoride as an additive in dental products. The most successful of these alternatives is sodium monofluorophosphate (MFP), which is now the most popular anti-decay additive in toothpastes. Some research suggests that another compound, stannous hexafluorozirconate, may be even more effective than either stannous fluoride or MFP as a decay preventive compound.

Stannous fluoride is a classic example of a compound that must be used in moderation. In small concentrations, it appears to be completely safe for human uses. In larger doses, it may have some health risks. For example, in concentrations of more than two parts per million, stannous fluoride (and other fluorides) may cause fluorosis, a mottling (brownish coloring) of the teeth and changes in bone composition. Individuals who work with pure stannous fluoride may also be at risk from exposure to its powder or dust. It is an irritant and may cause inflammation of the skin, eyes, or respiratory system. Symptoms may include coughing, wheezing, and shortness of breath. These effects do not occur, however, at the level at which stannous fluoride occurs in tooth care products.

FOR FURTHER INFORMATION

"How Crest Made Business History." Working Knowledge, Harvard Business School.
<http://hbswk.hbs.edu/item.jhtml?id=4574&t=bizhistory>
(accessed on November 15, 2005).

STANNOUS FLUORIDE

McCoy, Mike. "What's That Stuff: Fluoride."

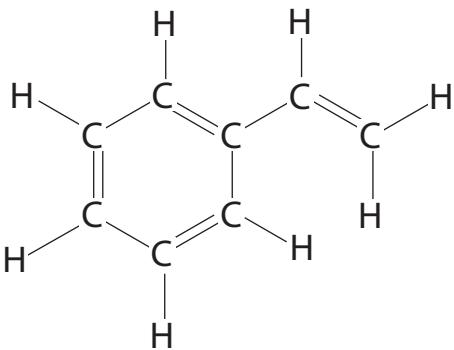
<http://pubs.acs.org/cen/whatstuff/stuff/7916sci4.html>
(accessed on November 15, 2005).

"Stannous Fluoride." International Programme for Chemical Safety.

<http://www.intox.org/databank/documents/chemical/stannousfluoride/eicso86o.htm> (accessed on November 15, 2005).

Walter, Patricia A. "Dental Hypersensitivity: A Review." *The Journal of Contemporary Dental Practice* (May 15, 2005): 107-117.

See Also Sodium Fluoride



OTHER NAMES:

Ethylbenzene; vinylbenzene; phenylethylene; styrol; styrolene; cinnamol

FORMULA:

C₆H₅CH=CH₂

ELEMENTS:

Carbon, hydrogen

COMPOUND TYPE:

Aromatic hydrocarbon (organic)

STATE:

Liquid

MOLECULAR WEIGHT:

104.15 g/mol

MELTING POINT:

-30.65°C (-23.17°F)

BOILING POINT:

145°C (293°F)

SOLUBILITY:

Insoluble in water; soluble in ethyl alcohol and acetone

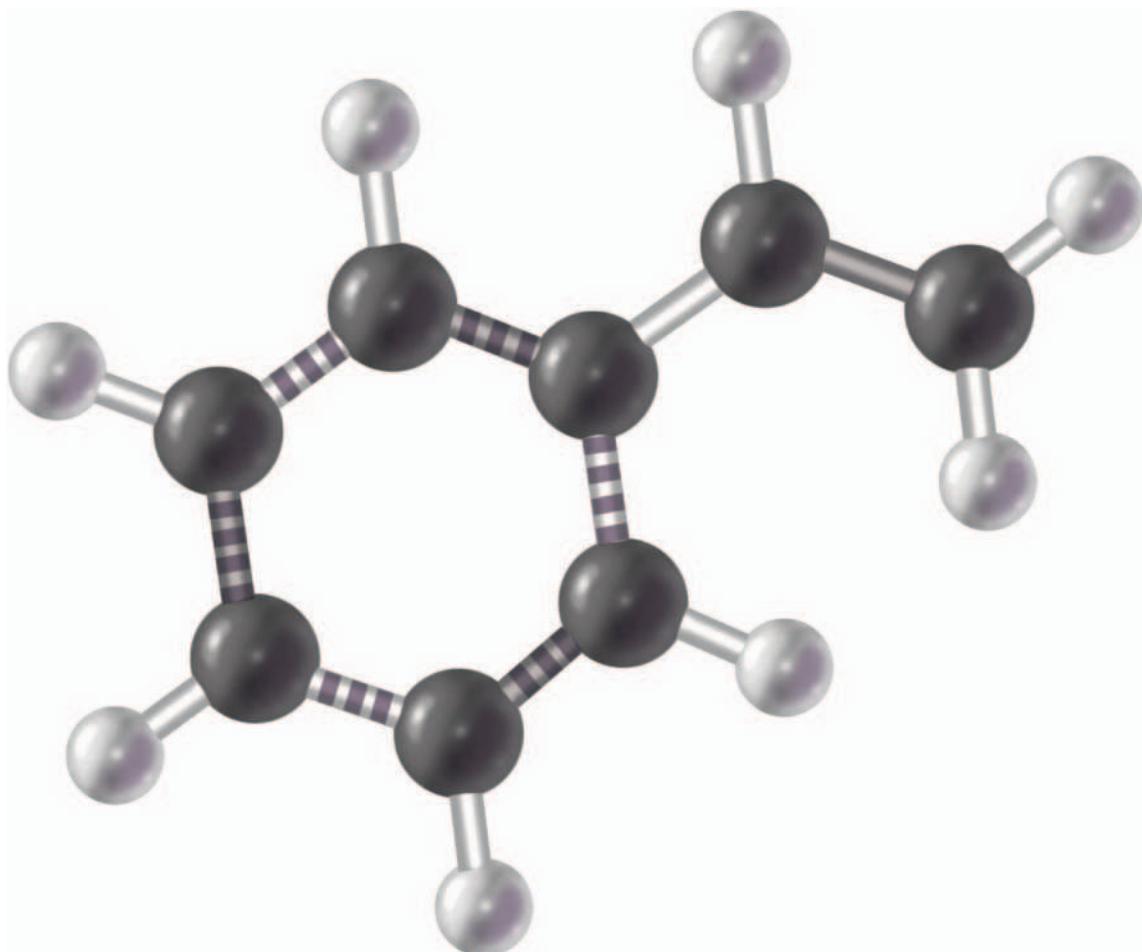
K E Y F A C T S

OVERVIEW

Styrene (STY-reen) is a colorless to yellowish, oily liquid with a sweet, flowery odor at low concentrations and a sharp, penetrating, disagreeable odor at high concentrations. When exposed to light and air, it slowly polymerizes. Polymerization is the chemical reaction in which a small molecule, such as styrene, reacts with other molecules of the same kind to produce very large molecules, usually made of long chains. Styrene occurs naturally in very small amounts in certain foods, such as coffee, strawberries, and cinnamon; in the sap of some trees; and in the gaseous emissions from internal combustion engines, waste incinerators, and tobacco smoke.

HOW IT IS MADE

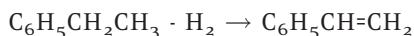
The first step in the preparation of styrene involves the reaction between benzene (C₆H₆) and ethylene (ethene; CH₂=CH₂), resulting in the formation of ethylbenzene (C₆H₅CH₂CH₃). The ethylbenzene is then dehydrogenated



Styrene. White atoms are hydrogen and black atoms are carbon. Gray stick indicates double bond; striped sticks show benzene ring.

PUBLISHERS RESOURCE GROUP

over a catalyst of iron(III) oxides at temperatures of about 600°C (1100°F). Dehydrogenation is the process by which hydrogen atoms are removed from a compound:



COMMON USES AND POTENTIAL HAZARDS

About two-thirds of all the styrene produced in the United States is used in the manufacture of polystyrene. Polystyrene is a clear, colorless, hard plastic that is easily molded and made into a foam known as styrofoam. It is used in the insulation of electrical wires and devices, in containers for hot and cold foods and drinks, and for the insulation of buildings.

Interesting Facts

- In 2004, about 5.4 million metric tons (5.9 million short tons) of styrene were produced in the United States. It ranked 17th among all chemicals produced that year.

Almost all of the remaining styrene is used in the production of other polymers, such as acrylonitrile-butadiene-styrene resins, styrene-acrylonitrile resins, styrene-butadiene rubber and latex, and various polyester resins.

The styrene fumes to which a person might be exposed are mild irritants to the skin, eyes, and respiratory tract. They may cause inflammation of the skin, eyes, nose, and throat. In larger amounts, they may adversely affect the nervous system causing nausea, tiredness, muscle weakness, depression, and concentration problems. Adequate studies on the carcinogenic properties of styrene have not been conducted, but the International Agency for Research on Cancer has called styrene a possible carcinogen. There are also no studies on the possible reproductive effects of exposure to styrene. In any case, the individuals most at risk for the health hazards of styrene are people who come into contact with the chemical in the workplace.

Styrene is also a moderate fire risk. In certain concentrations, it is also explosive. Again, these risks are of concern primarily to individuals who work with the pure compound in their jobs.

Words to Know

CARCINOGEN A substance that causes cancer in humans or other animals.

RESIN A natural or artificial soft solid material that is used as glue, ink, or in various plastics.

FOR FURTHER INFORMATION

“Material Safety Data Sheet: Styrene Monomer, Inhibited.” Department of Chemistry, Iowa State University.

<http://avogadro.chem.iastate.edu/MSDS/styrene.htm> (accessed on November 15, 2005).

“Styrene.” International Chemical Safety Cards.

<http://www.cdc.gov/niosh/ipcsneng/neng0073.html> (accessed on November 15, 2005).

“Styrene.” National Safety Council.

<http://www.nsc.org/library/chemical/styrene.htm> (accessed on November 15, 2005).

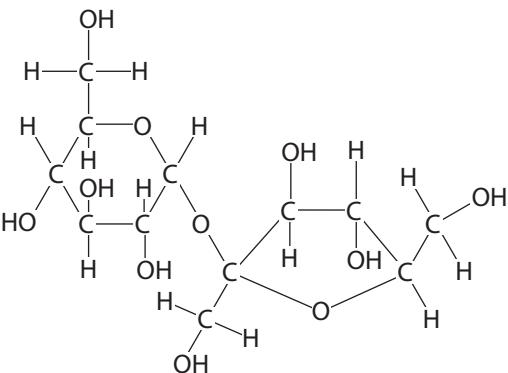
The Styrene Information and Research Center.

<http://www.styrene.org/> (accessed on November 15, 2005).

“ToxFAQs for Styrene.” Agency for Toxic Substances and Disease Registry.

<http://www.atsdr.cdc.gov/tfacts53.html> (accessed on November 15, 2005).

See Also Polystyrene



OTHER NAMES:

Saccharose; table sugar; beet sugar; cane sugar

FORMULA:

$C_{12}H_{22}O_{11}$

ELEMENTS:

Carbon, hydrogen, oxygen

COMPOUND TYPE:

Carbohydrate; disaccharide (organic)

STATE:

Solid

MOLECULAR WEIGHT:

342.30 g/mol

MELTING POINT:

185.5°C (365.9°F); begins to decompose at about 160°C (320°F)

BOILING POINT:

Not applicable

SOLUBILITY:

Soluble in water; slightly soluble in ethyl alcohol

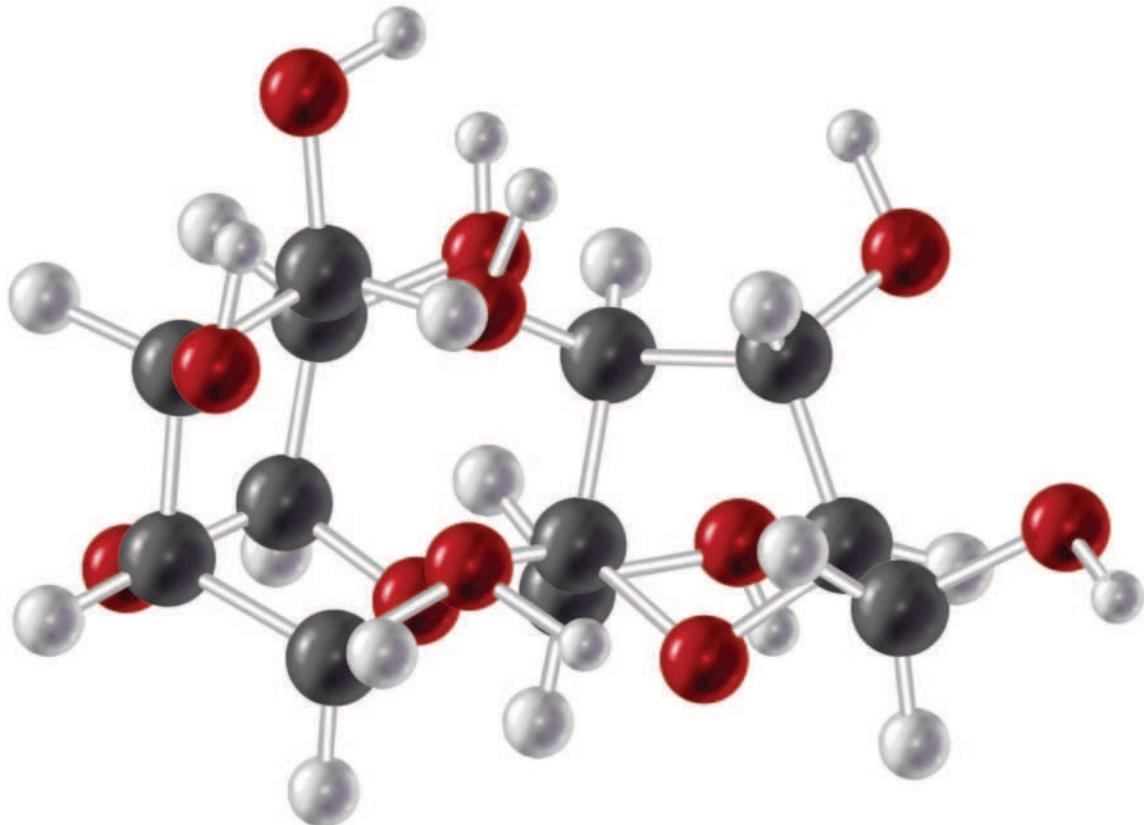
K
E
Y
F
A
C
T
S

OVERVIEW

Sucrose (SUE-kroze) is a white crystalline solid or powder with no odor but a characteristic sweet taste. It is the most widely used sweetener in the world. When heated, it tends to decompose, breaking down into carbon and water. The presence of carbon accounts for the increasing dark color of the compound as it caramelizes (changes from sucrose to caramel).

Sucrose is a disaccharide, a carbohydrate that consists of two monosaccharides. The carbohydrates are a large family of organic compounds that contain carbon, hydrogen, and oxygen, the latter two in a ratio of about two to one (as is the case with water). Thus the name “carbo-hydrate.” Some familiar monosaccharides are glucose, fructose, and galactose; while the disaccharides include sucrose, maltose (malt sugar), and lactose (milk sugar). Sucrose molecules consist of a molecule of glucose joined to a molecule of fructose.

The use of sugar is thought to have originated in Polynesia, from where it later spread to India. Probably the first



Sucrose. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. PUBLISHERS RESOURCE GROUP

recorded reference to sugar dates to 510 BCE when the Persian emperor Darius referred to sugar cane growing on the banks of the Indus River as “the reed which gives honey without bees.” Knowledge of sugar later spread across Asia into North Africa, but reached Europe only in the eleventh century. Warriors returning from the Crusades brought back stories about a new spice with a very pleasant flavor. Since growing sugar cane in Europe was difficult, the “new spice” was difficult to obtain and very expensive.

The sugar cane plant was brought to the New World on the earliest voyages of Christopher Columbus. His crew found that the plant grew easily in the congenial climate of the West Indies, and it rapidly became one of the most

important crops cultivated by Europeans. Indeed, some authors have observed that the profits from sugar exportation to Europe far exceeded the vast treasures of gold for which early explorers searched, and never found.

HOW IT IS MADE

Sucrose occurs naturally in a number of plants, primarily sugar cane and sugar beets. Sugar cane is the source of about 70 percent of the world's sucrose, and sugar beets account for the remaining 30 percent. Small amounts of sucrose are also obtained from sugar maple trees and sorghum, a cereal grain crop.

The process of making sugar depends on the source from which it is extracted. With sugar cane, the plant stalks are crushed to obtain the sweet juice they contain. The juice is treated with lime to remove impurities and then boiled until the juice thickens into a syrup. Evaporation of the syrup produces crystals of sucrose. Juice remaining after the evaporation step is sold as molasses. Stalks remaining after the juice has been extracted can be used as a fuel in the plant's boilers.

Sugar beets are harvested in the fields, washed, and cut into small pieces. The beet chips are then soaked in water and pressed to extract the sweet juice in them. The juice is boiled and the liquid evaporated to obtain crystalline sucrose. The solid material remaining from this process can be used as animal feed.

COMMON USES AND POTENTIAL HAZARDS

Sucrose is used almost entirely for human consumption in one form or another. The compound has two important functions. First, its sweet flavor makes foods more palatable and pleasant to eat. Second, its digestion results in the release of energy used by the body for growth, development, and everyday activities. When sucrose is digested, it first breaks down into the monosaccharides of which it is composed, glucose and fructose, with the release of energy stored in chemical bonds. Next, each of these monosaccharides is itself digested, resulting in the formation of carbon dioxide, water, and additional energy for body functions.

Sucrose is available in many forms, ranging from cubes to crystalline granules to finely ground sugar to powdered

Interesting Facts

- Early European physicians recommended sugar for the treatment of diseases. Both Buddha and Mohammad suggested its use as a medicine.
- The alcoholic drink rum was first produced in the West Indies in the early

sixteenth century by the distillation of sugar juice. Rum was first considered a drink suitable only for common people, but eventually spread to Europe and became popular among the upper classes as well.

sugar. All forms of sucrose have identical chemical and biological properties. Impure forms of sugar are also available. Brown sugar, for example, contains small amounts of molasses that are not removed during the sugar refining process.

Sucrose is used as a food additive in an almost unlimited variety of food products, including jams and jellies, candies, cakes and pies, ice cream and sherbet, and all types of canned and frozen foods. It serves a number of purposes in processed foods, including:

- To add sweetness to foods;
- As a preservative in jams and jellies;
- To increase the boiling point or reduce the freezing point of foods;
- To make possible the fermentation of foods by yeast;
- To improve the color or flavor of certain food products;
- As a demulcent, a substance that soothes the mucous membranes; and
- To add crispness to food products that do not have sufficient natural moisture.

Sucrose also has a limited number of industrial applications unrelated to the food industry. It is used in the production of some plastics, such as rigid polyurethane foams, in the manufacture of inks, and in the production of some types of soap.

Consumed in moderate amounts, sucrose poses no health hazards to humans or other animals. Eating excessive amounts of sucrose, however, is related to a number of health problems, most important, dental caries (tooth decay) and obesity. Sucrose is an important factor in the development of tooth decay because it provides the primary nutrient needed by bacteria living in the mouth. Cleaning one's teeth regularly is the best single way of preventing dental caries caused by the consumption of sucrose. Obesity is a problem that develops when a person consumes more foods than needed for normal healthy body function. When those foods are not utilized by the body, they are stored as fatty deposits, resulting in obesity and a number of health problems related to it, such as diabetes and heart disease.

FOR FURTHER INFORMATION

Mardis, Anne L. "Current Knowledge of the Health Effects of Sugar Intake." *Family Economics and Nutrition Review* (Winter 2001): 88-91. Also available online at <http://www.cnpp.usda.gov/FENR/FENRv13n1/fenrv13n1p87.pdf>.

"Sucrose." J. T. Baker.
<http://www.jtbaker.com/msds/englishhtml/s7394.htm>
(accessed on November 15, 2005).

"Sugar. Sweet by Nature." The Sugar Association.
<http://www.sugar.org/> (accessed on November 15, 2005).

"Welcome to the World of Sugar Technology." Sugar Knowledge International.
<http://www.sucrose.com/> (accessed on November 15, 2005).

"What Is Sugar?" The Accidental Scientist.
<http://www.exploratorium.edu/cooking/candy/sugar.html>
(accessed on November 15, 2005).

See Also Fructose; Glucose; Lactose; Sucrose Polyester

OTHER NAMES:	Sucrose octaoleate; sucrose octatester
FORMULA:	See Overview
ELEMENTS:	Carbon, hydrogen, oxygen
COMPOUND TYPE:	Polyester (organic)
STATE:	Liquid or solid
MOLECULAR WEIGHT:	Varies
MELTING POINT:	Not applicable
BOILING POINT:	Not applicable
SOLUBILITY:	Insoluble in water; soluble in many organic solvents

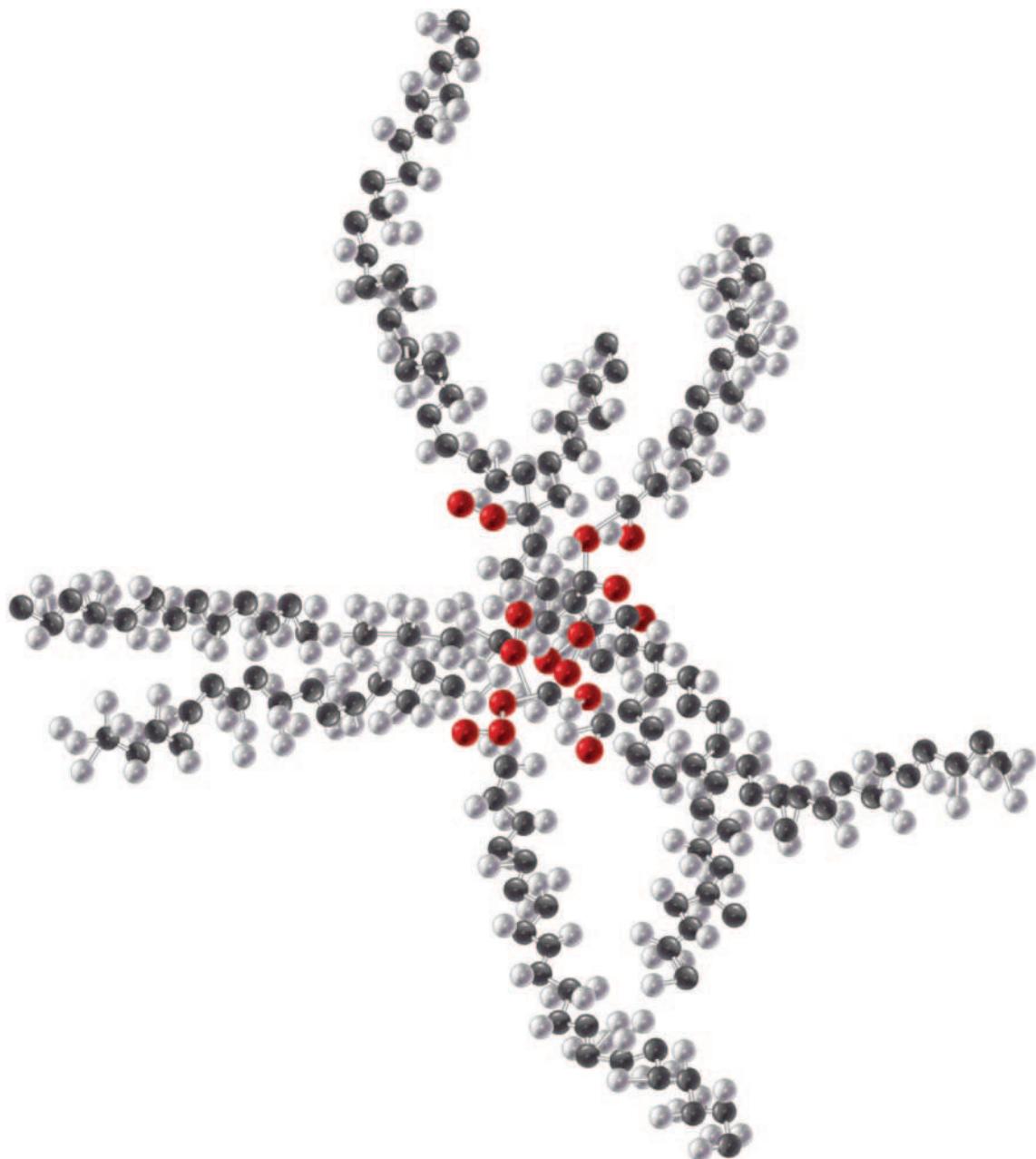
KEY FACTS

Sucrose Polyester

OVERVIEW

Sucrose polyester (SUE-kroze pol-ee-ESS-ter) is an artificial fat available under the trade names of Olestra® or Olean®. The sucrose polyester molecule is quite large and cannot be absorbed or digested by the human digestive system. Since it cannot be digested, it provides no calories to one's diet. The compound is used primarily as an additive in snack treats targeted at people on a diet.

Sucrose polyester is one of many compounds developed by researchers in an effort to reduce the caloric intake of people trying to lose weight. Artificial sweeteners designed for this purpose include saccharin, cyclamates, aspartame, and acesulfame K. Sucrose polyester is probably the best known and most successful artificial fat produced so far. It was developed in the late 1960s by two researchers at Procter & Gamble's Miami Valley Laboratories, Robert Volpenhein and Fred Mattson. Volpenhein and Mattson were looking for a new kind of fat that could be digested more easily by premature babies. Instead, they made a very different



Sucrose octaoleate. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. PUBLISHERS RESOURCE GROUP

discovery: a compound that has many of the properties of a fat (including its flavor), but that passes through the digestive system without being absorbed or digested. Company executives immediately saw the practical application of the

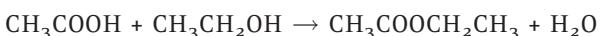
new compound for dieters and began a program for gaining federal approval for its use as a food additive.

The regulatory history of sucrose polyester has been very complicated, however. It was not until 1987 that Procter & Gamble had gained enough research evidence to submit sucrose polyester to the U.S. Food and Drug Administration (FDA) for its approval. Then, it took another ten years for the FDA to act on (and approve) Procter & Gamble's application. Even then, FDA approval was limited to certain specialized kinds of foods, including potato chips, crackers, and tortilla chips. The FDA's action has not been repeated in many other countries, however. For example, the United Kingdom has never approved the use of sucrose polyester as a food additive in that nation.

In fact, Procter & Gamble has faced a number of obstacles in marketing products made with sucrose polyester. A number of questions have been raised about possible health effects of the compound, and very few products containing sucrose polyester are currently available in the marketplace.

HOW IT IS MADE

Sucrose is a molecule that has eight hydroxyl (-OH) groups in it. For that reason, it can be thought of as a *polyol*. A polyol is an alcohol with two or more hydroxyl groups. A characteristic reaction of alcohols is that they react with acids to form esters. As a simple example, the reaction between acetic acid and ethanol produces the ester ethyl acetate:



Perhaps you can imagine what a molecule of sucrose polyester looks like. It has the basic structure of sucrose

Interesting Facts

- The first commercial product in the United States to contain Olestra was the Frito Lay WOW brand of potato chips.
- One ounce of potato chips made with Olestra has 75 calories and no fat; by comparison, an ounce of regular potato chips have 150 calories and 10 grams of fat.

with six to eight fatty acids attached to it. The molecule is very large and cumbersome. No wonder the digestive system is unable to break it down!

A number of methods are available for reacting all or most of the hydroxyl groups in sucrose with fatty acids. In one method, the hydroxyl groups are first reacted with a much simpler acid, acetic acid (CH_3COOH), forming sucrose octaacetate (*octa* = “eight”). The sucrose octaacetate is then reacted with a compound that has fatty acid fragments in it, such as methyl palmitate. The sucrose octaacetate is converted to sucrose polyester.

COMMON USES AND POTENTIAL HAZARDS

Procter & Gamble had high hopes for Olestra when it was first introduced to the marketplace in the late 1990s. Public reaction was very positive at first. However, enthusiasm for the new product began to die off rather quickly. Olestra-containing products earned the company \$400 million in 1998, but only two years later, sales had dropped to \$200 million. One concern is that Olestra may have undesirable health effects on some individuals. These effects include bloating, diarrhea, cramps, loose stools, and urgency of defecation. The compound also appears to interfere with the absorption of certain vitamins. When the FDA approved Olestra for use in the United States, it added the requirement that vitamins A, D, E, and K also be added to counteract this effect. The future status of sucrose polyester is, at this point, somewhat in doubt.

FOR FURTHER INFORMATION

"A Brief History of Olestra." Center for Science in the Public Interest.

<http://www.cspinet.org/olestra/history.html> (accessed on November 15, 2005).

"Fat Replacers: Olestra." College of Agricultural Sciences, Pennsylvania State University.

<http://pubs.cas.psu.edu/FreePubs/pdfs/uko58.pdf> (accessed on December 29, 2005).

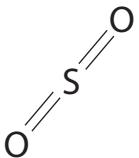
"Pass the Potato Chips!" Vanderbilt University.

http://www.vanderbilt.edu/AnS/psychology/health_psychology/olestra.htm (accessed on November 15, 2005).

"Welcome to Olean Brand Olestra." Procter & Gamble.

<http://www.olean.com/> (accessed on November 15, 2005).

See Also Sucrose



OTHER NAMES:	Sulfurous oxide; sulfurous anhydride
FORMULA:	SO_2
ELEMENTS:	Sulfur, oxygen
COMPOUND TYPE:	Nonmetallic oxide (inorganic)
STATE:	Gas
MOLECULAR WEIGHT:	64.06 g/mol
MELTING POINT:	-75.5°C (-104°F)
BOILING POINT:	-10.05°C (-13.91°F)
SOLUBILITY:	Soluble in water, ethyl alcohol, ether, and chloroform

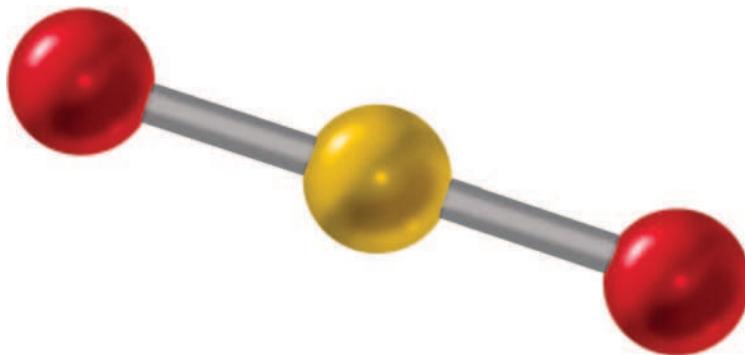
KEY FACTS

Sulfur Dioxide

OVERVIEW

Sulfur dioxide (SUL-fur dye-OK-side) is a colorless gas with a sharp, harsh odor similar to that of a burning match. It can act as both an oxidizing agent and a reducing agent. An oxidizing agent is a substance that provides oxygen to other substances or provides electrons to them. A reducing agent removes oxygen from other substances or removes electrons from them. Sulfur dioxide dissolves readily in water, forming sulfurous acid (H_2SO_3), which is readily converted to sulfuric acid (H_2SO_4).

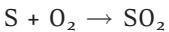
Sulfur dioxide is a natural component of air. It is produced when trees, brush, and other organic matter burn, and it is present in volcanic gases. It is also released during the normal metabolic reactions of some living organisms, especially marine plankton and bacteria. The compound is also produced in large amounts by human activities. All fossil fuels (coal, oil, and natural gas) contain small amounts of sulfur as impurities. When those fuels burn, the sulfur they contain is converted to sulfur dioxide, which becomes a component of air pollution.



Sulfur dioxide. Red atoms are oxygen and yellow atom is sulfur. PUBLISHERS RESOURCE GROUP

HOW IT IS MADE

Sulfur dioxide can be prepared by several methods, the most common of which is the combustion of sulfur or pyrites (FeS_2). A variety of furnaces have been developed for carrying out this reaction. Each type of furnace produces sulfur dioxide of different purities. After production, the sulfur dioxide is normally cooled and compressed to convert it to liquid form. Liquid sulfur dioxide is more easily stored and transported than the gaseous form. Sulfur dioxide is also obtained as the byproduct of a number of industrial operations, especially the smelting of metallic ores. Smelting is the process by which a metal is extracted from its ore by heating in air. Since many ores are sulfides, this process often results in the formation of sulfur dioxide, which can be captured as a byproduct of the operation. Finally, sulfur dioxide can be produced by the direct combustion of sulfur itself:



COMMON USES AND POTENTIAL HAZARDS

About 45 percent of all the sulfur dioxide produced in the United States is used in the manufacture of other chemical compounds, the most important of which is sodium bisulfite (NaHSO_3). Other compounds made from sulfur dioxide include sulfuric acid (H_2SO_4), chlorine dioxide (ClO_2), sodium dithionite ($\text{Na}_2\text{S}_2\text{O}_6 \cdot 2\text{H}_2\text{O}$), and sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$). Sulfur dioxide is also used as a bleaching agent for a number of products, including pulp and paper, textile fibers, straw, glue, gelatin, starches, grains, and various oils. The compound

Interesting Facts

- Sulfur dioxide was first studied in detail by the English physicist and chemist Joseph Priestley (1733-1804), who invented a method for collecting gases over water.
- The ancient Greeks and Romans fumigated their homes by burning sulfur. The sulfur dioxide formed destroyed microorganisms that cause disease and rot.
- The concentration of sulfur dioxide in clean air above the continents is less than one part per billion. Volcanic eruptions account for about half of all the gas produced by natural sources.

has a number of agricultural uses, especially in the treatment of soybeans and corn to destroy molds and preserve the product from decay. It is also used in the processing and refining of metal ores and petroleum. For example, it is added to some petroleum products to remove dissolved oxygen that would cause rust of pipes through which the products are distributed. Some other applications of the compound include:

- As a preservative for certain dried fruits and vegetables, such as cherries and apricots;
- As an additive in beers to prevent the formation of harmful products known as nitrosamines;
- As an additive in wines to prevent the growth of undesirable molds and other fungi;
- In the production of high-fructose corn syrups (HFCS) used as sweeteners in commercial food and drink products;
- As an antichlor in water purification systems;
- In the refining of sugar;
- In the manufacture of certain clay products to counteract the presence of compounds or iron and other metals that would impart color to the final product;

- In the molding and casting of magnesium parts and products; and
- In the sulfonation of oils, an important chemical process by which a sulfate group (-SO_3) is added to a compound.

Sulfur dioxide is a very toxic gas that can be an irritant to the eyes, the respiratory system, and, in some cases, the skin. At concentrations normally found in ambient air (the typical atmospheric environment surrounding us), these effects are annoying, but not particularly dangerous. Such is not the case for individuals with respiratory disorders, the young, or the elderly. Such individuals may experience more serious breathing problems that require medical attention. Higher concentrations of the gas may cause more serious problems, such as coughing, headache, dizziness, feelings of suffocation, and nausea. These conditions are most likely to occur in areas where air pollution is a problem, as in urban or industrial areas. People who are constantly exposed to relatively high concentrations of sulfur dioxide (such as smelter workers) may experience more serious long-term health problems, such as asthma, chronic bronchitis, lung disease, or emphysema.

The health effects of sulfur dioxide are serious enough that the compound is not allowed as a food additive in meats and other food products that contain vitamin B₁. The reason is that sulfur dioxide reacts with and destroys the vitamin. The U.S. Environmental Protection Agency (EPA) has determined that humans should not be exposed to a concentration of more than 0.030 parts per million on average throughout the year, or more than 0.14 parts per million over any one 24-hour period.

In addition to its health effects on humans, sulfur dioxide has some important consequences for the physical and biological environment. Those effects occur because sulfur dioxide released to the atmosphere from electricity-generating plants and factories combines with moisture in the air to form sulfuric acid. Sulfuric acid then falls to earth in the form of acid rain, acid snow, or some other form of acid precipitation where it damages buildings and other structures, trees and other plant life, and fish and other aquatic organisms. Since 1995, the EPA has sponsored a variety of control programs designed to reduce the release of sulfur dioxide into the atmosphere to prevent such problems.

Words to Know

ANTICHLOR A chemical that reacts with excess chlorine used for purification, disinfecting, or some other purpose.

METABOLISM All of the chemical reactions that occur in cells by which fats, carbohy-

drates, and other compounds are broken down to produce energy and the compounds needed to build new cells and tissues.

FOR FURTHER INFORMATION

“Acid Rain Program.” U.S. Environmental Protection Agency.
<http://www.epa.gov/airmarkets/arp/> (accessed on November 15, 2005).

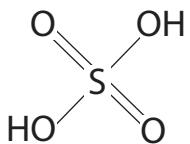
“SO₂, How Sulfur Dioxide Affects the Way We Live & Breathe.” Office of Air Quality Planning and Standards, U.S. Environmental Protection Agency. November 2000. Also available online at
<http://www.epa.gov/air/urbanair/so2>. (accessed on November 15, 2005).

“Sulfur Dioxide.” Air Liquide.
http://www.airliquide.com/en/business/products/gases/gas_data/index.asp?GasID=27 (accessed on November 15, 2005).

“Sulfur Dioxide.” New Jersey Department of Health and Senior Services.
<http://www.state.nj.us/health/eoh/rtkweb/1759.pdf> (accessed on November 15, 2005).

“ToxFAQs for Sulfur Dioxide.” Agency for Toxic Substances and Disease Registry.
<http://www.atsdr.cdc.gov/tfacts116.html> (accessed on November 15, 2005).

See Also Sodium Sulfite; Sulfuric Acid



OTHER NAMES:	Hydrogen sulfate; oil of vitriol
FORMULA:	H_2SO_4
ELEMENTS:	Hydrogen, sulfur, oxygen
COMPOUND TYPE:	Acid (inorganic)
STATE:	Liquid
MOLECULAR WEIGHT:	98.08 g/mol
MELTING POINT:	10.31°C (50.56°F)
BOILING POINT:	337°C (639°F)
SOLUBILITY:	Completely miscible with water

KEY FACTS

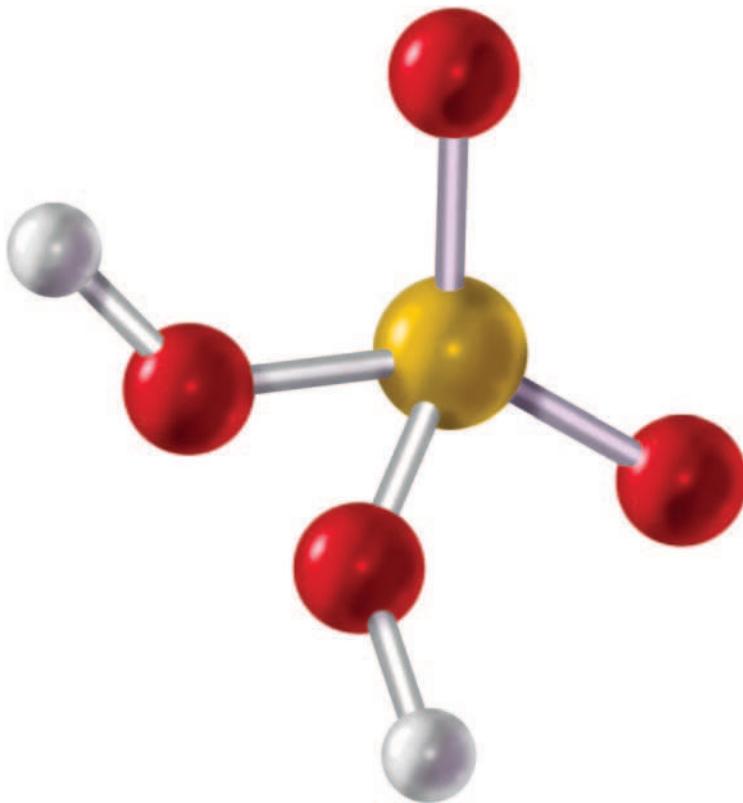
Sulfuric Acid

OVERVIEW

Sulfuric acid (sul-FUR-ik AS-id) is a colorless to dark brown, dense, oily liquid that mixes completely with water. Its color depends on its purity, with compounds of sulfur adding a darker color to the colorless pure product. Sulfuric acid is also available in a form known as fuming sulfuric acid, a solution of sulfur trioxide in sulfuric acid with the chemical formula $x\text{H}_2\text{SO}_4 \cdot y\text{SO}_3$. Sulfuric acid is the most commonly manufactured chemical in the world. In 2004, the chemical industry produced 37,515,000 metric tons (41,266,000 short tons), of which more than half was used in the manufacture of fertilizers.

Some historians credit the discovery of sulfuric acid to the Islamic scientist Mohammad Ibn Zakariya al-Razi (864-930 CE), while others claim the first mention of the compound to have been by the Islamic writer Geber (probably born about 1270). In any case, European scientists apparently did not discover sulfuric acid on their own until the sixteenth century, when the Belgian scientist Johann Baptista

Sulfuric acid. Red atoms are oxygen; white atoms are hydrogen; and yellow atom is sulfur. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

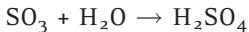
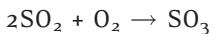
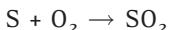


van Helmont (1579-1644) described its preparation by adding water to the gas formed when sulfur was burned. The discovery of sulfuric acid proved to be an important development in the early history of modern chemistry. For the first time, it gave chemists an acid far stronger than the vinegar with which they previously had to work in analyzing mixtures and making new chemical compounds.

HOW IT IS MADE

The first commercially successful method for making sulfuric acid was developed in 1746 by English physician, chemist, and inventor John Roebuck (1718-1794). Roebuck's method is called the *lead chamber process* because the acid is made in large containers lined with lead. The lead chamber process involves three primary steps: the combustion of sulfur to produce sulfur dioxide; the conversion of sulfur

dioxide to sulfur trioxide; and the reaction of sulfur trioxide with water to make sulfuric acid:



Although the process looks fairly simple, the reactions are actually somewhat complex because of other compounds needed to make the reactions occur.

The second method for making sulfuric acid is known as the *contact process*. It was invented about 1830 by an English vinegar merchant from Bristol, Peregrine Phillips. The chemical reactions involved in Phillips' process are identical to those in the lead chamber process, but they are carried out over a catalyst of finely divided platinum metal. Phillips found that the yield obtained (the amount of raw material converted to useful product) was much higher than with the lead chamber process.

As it happens, little attention was paid to Phillips' discovery because there was not much demand for sulfuric acid at the time. It was not until the invention of synthetic dyes a few decades later that the compound became commercially important. But even then, the lead chamber process was the preferred method for making sulfuric acid. Over time, improvements were made in the contact process, and it gradually became more and more popular. Today, nearly all of the sulfuric acid produced is manufactured by some modification of Peregrine Phillips' method.

COMMON USES AND POTENTIAL HAZARDS

The most important single use of sulfuric acid is for the production of phosphoric acid (H_3PO_4), which in turn is used to make fertilizers. About 70 percent of all the sulfuric acid used in the United States goes to this application. Some other uses of sulfuric acid include:

- As the electrolyte in lead storage batteries (the liquid through which charged particles flow);
- For the processing of gasoline and other petrochemicals, as a way of removing impurities present in the products;

Interesting Facts

- Sulfuric acid's common name is oil of vitriol. That name comes from the fact that it was once produced from either iron(II) sulfate, known as "green vitriol," or from copper(II) sulfate, known as "blue vitriol."
- The conversion of sulfur dioxide to sulfur trioxide in the manufacture of sulfuric acid does not occur very readily. That step controls

the efficiency with which the acid can be made. Phillips' great breakthrough was to find a catalyst that made that reaction take place more easily. Later, researchers found an even more efficient catalyst, vanadium pentoxide (V_2O_5). Today, vanadium pentoxide is still the most popular catalyst in the production of sulfuric acid.

- As a cleaning agent for metal surfaces, especially prior to their being plated with a second metal;
- In the manufacture of explosives, dyes, glues, pigments, rayon, and films;
- As a catalyst in a variety of industrial chemical and research chemical reactions; and
- In the processing of ores in preparation for the extraction of metals.

Sulfuric acid is a highly corrosive material that causes severe damage to the skin, eyes, and respiratory system. If spilled on the skin or eyes, it can cause inflammation, burns, and blistering, and cause serious damage to one's vision. If

Words to Know

CATALYST A material that increases the rate of a chemical reaction without undergoing any change in its own chemical structure.

MISCIBLE Able to be mixed; especially applies to the mixing of one liquid with another.

swallowed, it causes severe burns of the gastrointestinal tract and irreversible damage to tissue and organs. Any accident in which sulfuric acid is inhaled, swallowed, or spilled on the body requires immediate medical attention.

FOR FURTHER INFORMATION

“Chemical of the Week: Sulfuric Acid.” Science is Fun.

<http://scifun.chem.wisc.edu/chemweek/Sulf&Top/Sulf&Top.html> (accessed on November 15, 2005).

“Chronic Toxicity Summary: Sulfuric Acid.” Office of Environmental Health Hazard, State of California.

http://www.oehha.org/air/chronic_rels/pdf/sulfuric.pdf (accessed on November 15, 2005).

“Sulfuric Acid.” In *World of Scientific Discovery*. 2nd ed. Detroit: Gale Group, 1999.

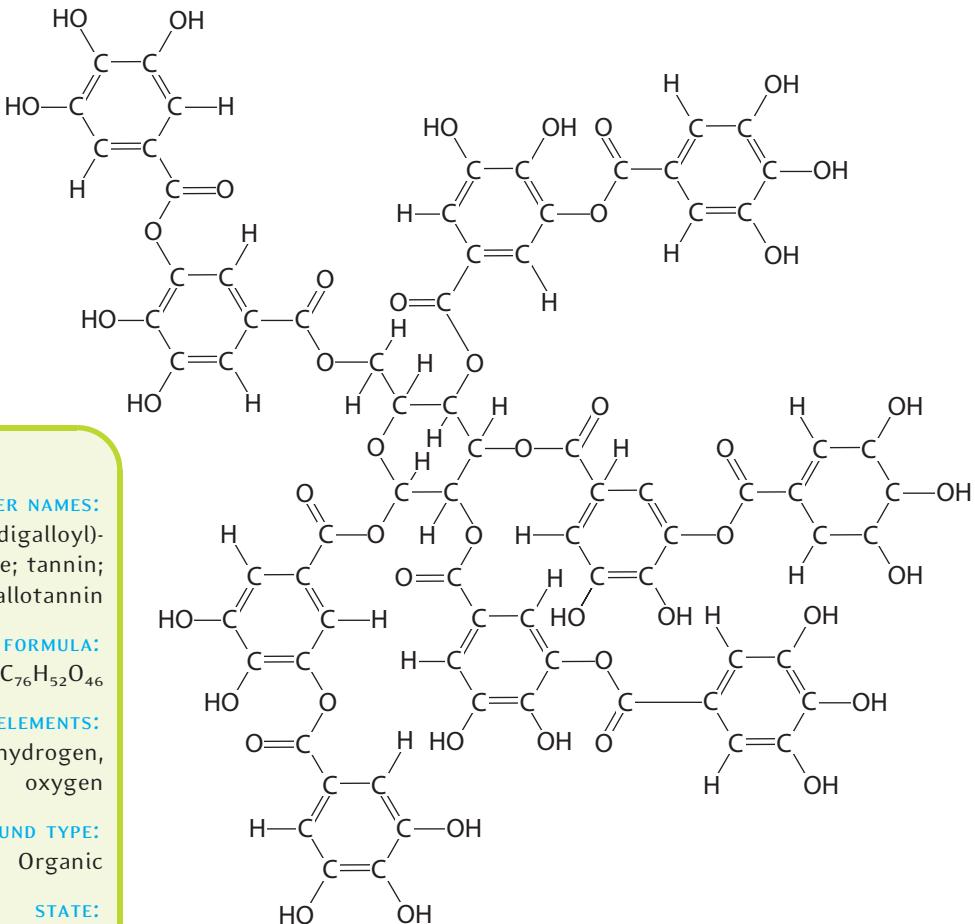
“Sulfuric Acid.” National Safety Council.

<http://www.nsc.org/library/chemical/sulfuric.htm> (accessed on November 15, 2005).

“Sulfuric Acid: Production and Uses.” Aus e tute.

<http://www.ausetute.com.au/sulfacid.html> (accessed on November 15, 2005).

See Also Phosphoric Acid; Sulfur Dioxide



OTHER NAMES:
Penta-(m-digalloyl)-glucose; tannin; gallotannin

FORMULA:
 $C_{76}H_{52}O_{46}$

ELEMENTS:
Carbon, hydrogen, oxygen

COMPOUND TYPE:
Organic

STATE:
Solid

MOLECULAR WEIGHT:
1701.20 g/mol

MELTING POINT:
Begins to decompose at 210°C (410°F)

BOILING POINT:
Not applicable

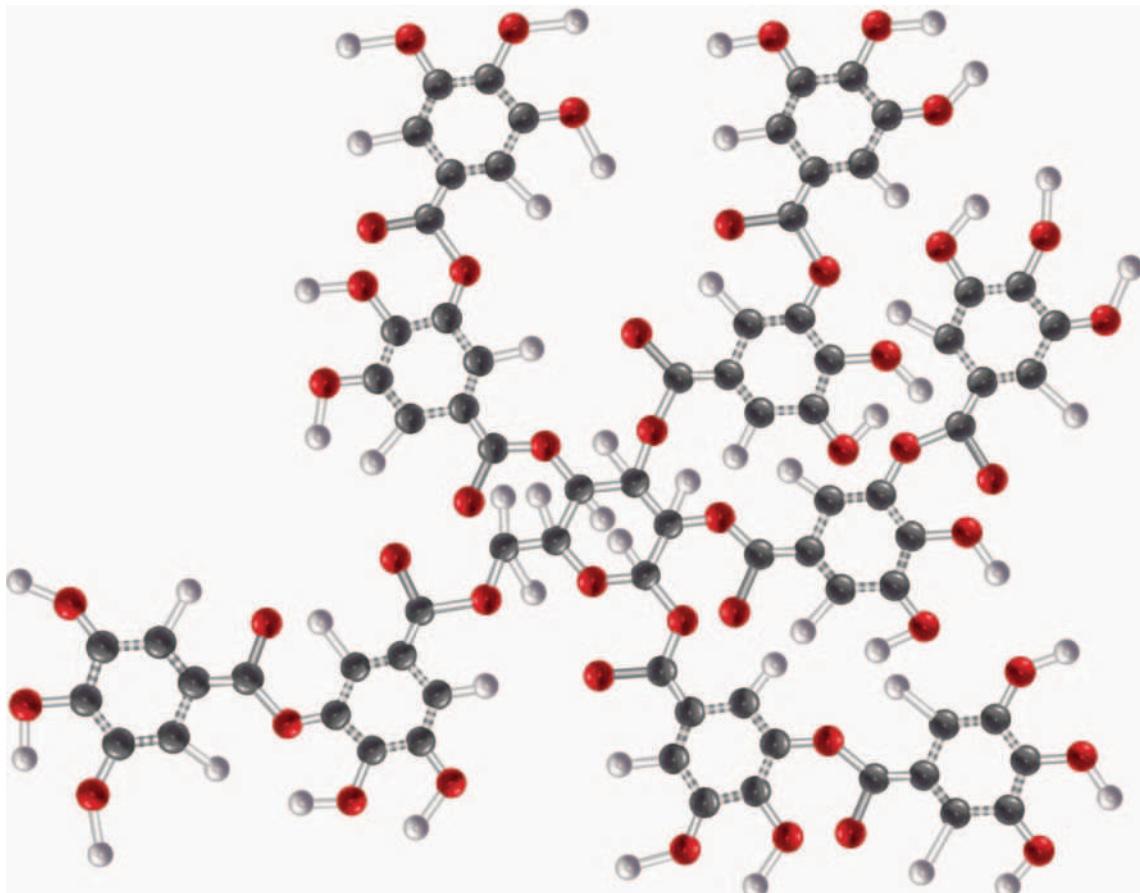
SOLUBILITY:
Soluble in water, ethyl alcohol, and acetone

KEY FACTS

Tannic Acid

OVERVIEW

Tannic acid (TAN-ik AS-id) is a pale yellow amorphous powder, shiny scales, or spongy material that gradually darkens when exposed to air. It is odorless but has a strong, bitter taste. It is a derivative of the simple sugar glucose ($C_6H_{12}O_6$) in which five hydroxyl groups (-OH) have been replaced by large complex side chains known as digalloyl groups. Tannic acid occurs naturally in the bark of hemlock, chestnut, mangrove, and oak trees; in sumac plants; and in



Tannic acid. Red atoms are oxygen; black atoms are carbon; and white atoms are hydrogen.

PUBLISHERS RESOURCE GROUP

plant galls, the hard swellings that develop on trees infested with insect parasites. The name *tannic acid* comes from the fact that the compound is used in the process known as tanning. It transforms the proteins in raw animal hides into forms that resist the natural process of decomposition, converting the hides into leather. Tannic acid belongs to a large class of chemical compounds known as tannins that have the common chemical property of being able to precipitate proteins. Precipitation is the process by which the soluble form of a substance is converted to an insoluble form, which then settles out of solution. In addition to its applications in tanning, tannic acid is also used for staining wood, dyeing cloth, and treating minor cuts and wounds.

Interesting Facts

- A kind of ink called *iron gall ink* is made from tannic acid, iron(II) sulfate, gum arabic, and water. The recipe for this ink has been known since the Middle Ages. It is used for important documents because it does not wear away easily.
- Tannic acid is mentioned in some very old medical books as a treatment for mushroom poisoning.
- Studies suggest that people allergic to cats may be able to control their allergies by spraying surfaces in the home with a solution of tannic acid. The compound chemically alters the proteins in feline saliva that trigger allergic reactions.
- Tannic acid is not a true acid because it does not contain the carboxyl group (-COOH) present in all organic acids. Instead, it is a polyphenol, a compound with many ("poly") phenol (C_6H_5OH) groups.

HOW IT IS MADE

Tannic acid is extracted from natural source by using the principle that the compound is soluble in water and alcohol, while many components with which it occurs are not. The most common sources of tannic acid have traditionally been the Chinese nutgall, which consists of about 70 percent tannic acid, and Turkish nutgall, which contains about 50 to 60 percent tannic acid. The nutgall is ground into a powder and then mixed with a combination of water and/or alcohol and ether. The tannic acid in the powder dissolves in the water/alcohol portion of the mixture, and other components of the gall dissolve in the ether portion. The water and alcohol are then allowed to evaporate, leaving behind relatively pure tannic acid. In recent years, tannic acid has been obtained more economically from the bark of the quebracho tree, which grows in South America and is related to poison ivy.

COMMON USES AND POTENTIAL HAZARDS

Tannic acid continues to be used for many of the applications for which it has been known for centuries, including

Words to Know

AMORPHOUS Without crystalline structure.

ELECTROPLATING A process by which a thin layer of one metal is deposited on top of a second metal by passing an electric current through a solution of the first metal.

MORDANT A substance used in dyeing and printing that reacts chemically with both a dye and the material being dyed to

help hold the dye permanently to the material.

MUTAGEN A substance that causes a mutation in plants or animals. Mutations are changes in an organism's genetic composition.

PRECIPITATE A solid material that settles out of a solution, often as the result of a chemical reaction.

the tanning of animal hides, the manufacture of specialized inks, and the treatment of minor cuts and bruises. Its medical applications are based on its astringent properties. An astringent is a compound that triggers a loss of water from tissue, thereby causing the tissues to shrink and contract. The compound also has a number of other commercial and industrial uses, including:

- As a mordant in the coloring of fabrics and the printing of colored papers;
- In the manufacture of many kinds of chemicals, including tannates (compounds of tannic acid and a metal), gallic acid ($C_6H_2(OH)_3COOH$; with many of the same uses as tannic acid), and pyrogallic acid ($C_6H_3(OH)_3$; also with similar uses);
- As a clarifying (purifying) agent in the manufacture of beers and wines;
- In electroplating;
- In the manufacture of artificial horn and tortoise shell products;
- As a coagulant in the manufacture of rubber to precipitate out impurities;
- In the deodorizing of crude oil; and
- In photographic processes.

Tannic acid is a mild irritant of the skin, eyes, and respiratory tract. It is toxic by ingestion. In large doses, it may cause liver damage. No evidence is available on its carcinogenic or mutagenic effects in humans.

FOR FURTHER INFORMATION

Eusman, Elmer. "The Ink Corrosion Website."

<http://www.knaw.nl/ecpa/ink/> (accessed on November 15, 2005).

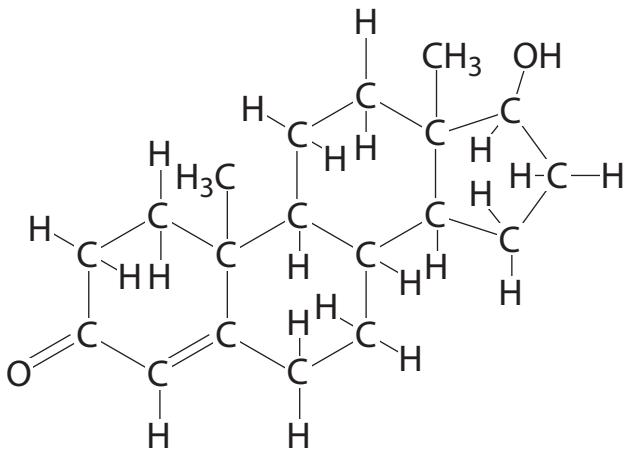
"Material Safety Data Sheet: Tannic Acid." GFS Chemicals. <https://gfschemicals.com/Search/MSDS/2415MSDS.PDF> (accessed on November 15, 2005).

Meyer, John R. "By Gall y."

<http://www.cals.ncsu.edu/course/ent591k/gally.html>
(accessed on November 15, 2005).

"Tannic Acid/Tannins." J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/too65.htm>
(accessed on November 15, 2005).

**OTHER NAMES:**

See Overview

FORMULA: $C_{19}H_{28}O_2$ **ELEMENTS:**Carbon, hydrogen,
oxygen**COMPOUND TYPE:**

Steroid (organic)

STATE:

Solid

MOLECULAR WEIGHT:

288.41 g/mol

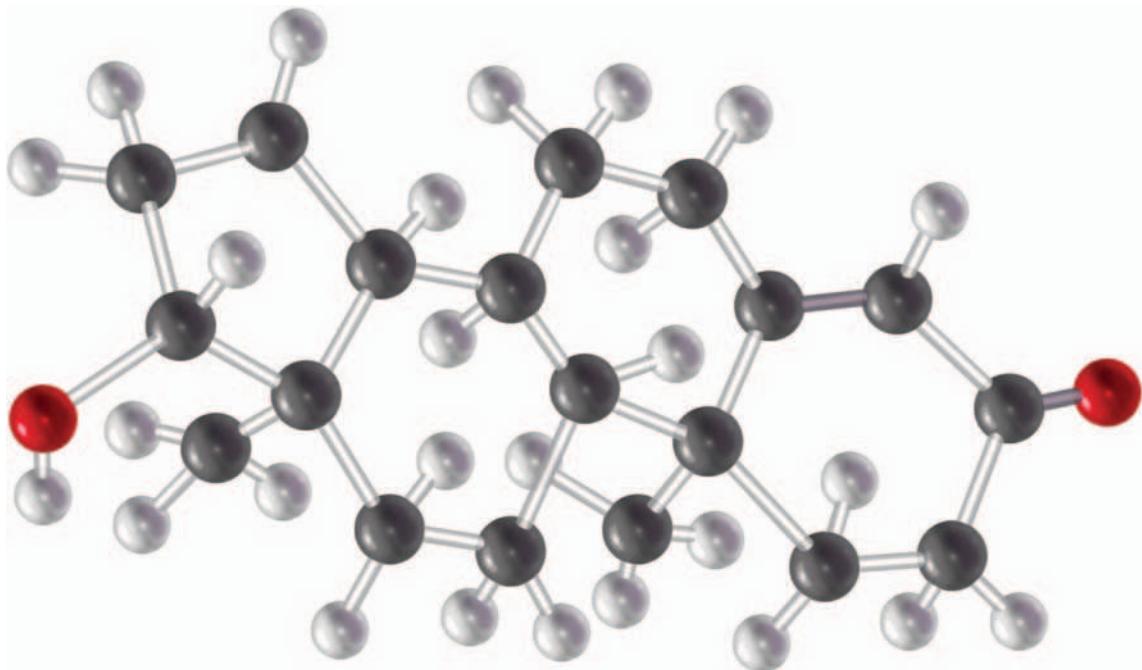
MELTING POINT:

155°C (311°F)

BOILING POINT:Decomposes above
melting point**SOLUBILITY:**Insoluble in water;
soluble in ethyl
alcohol, chloroform,
vegetable oils, ether,
and other organic
solvents**K E Y F A C T S****OVERVIEW**

Testosterone (tes-TOS-ter-own) is a white to creamy-white crystalline or powdery material that is odorless and tasteless and stable in air. It is a member of the organic family of compounds known as androgenic steroids—hormones responsible for the development of male sexual characteristics, such as male sex organs, a deep voice, and facial hair. Testosterone is present in both male and female bodies, although to a greater extent in men than in women.

The earliest studies of testosterone were conducted by French physiologist Charles E. Brown-Séquard (1817–94), sometimes called the father of endocrinology. Endocrinology is the study of hormones, their effects, and the glands that produce them. In 1889, at the age of 72, Brown-Séquard injected himself with extracts from dog and pig testes and reported a major change in his physical and mental health. He reported that he had regained the strength that he had as a younger man. He also said the extract improved his intellectual abilities. Brown-Séquard's scientific colleagues



Testosterone. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

laughed off his discoveries, and studies of testosterone were largely suspended for almost half a century.

In the 1930s, however, interest in the compound was revived. In 1935, scientists obtained the first pure sample from testosterone and were able to confirm Brown-Séquard's findings. Within a matter of years, the compound became especially popular among middle-aged men who believed that it could restore their physical and mental abilities. Testosterone was first synthesized in 1935 by German chemist Adolf Friedrich Johann Butenandt (1903–1995), an accomplishment for which he was awarded the 1939 Nobel Prize in Chemistry.

Testosterone is also known by the following names; 17β -Hydroxyandrost-4-ene-3-one; androstan-4-en- 17β -ol-3-one; testosterone; testosteron; testostosterone; and trans-testosterone.

HOW IT IS MADE

Testosterone is produced naturally in the male testes and the female ovaries. It is also made synthetically starting either with cholesterol or diosgenin, a plant steroid.

Interesting Facts

- Women who choose to undergo a series of procedures to change their gender are required to take testosterone over a period of many months to stimulate the development of male sex characteristics. They must continue taking testosterone supplements for the rest of their lives.

COMMON USES AND POTENTIAL HAZARDS

Both males and females of all vertebrate species produce testosterone. The amount present in the male body is considerably greater than that present in the female body. Testosterone has a number of biological effects on the body, including an increase in the number of red blood cells and muscles cells and initiation of the development of male sex organs. It is also responsible for the development of secondary male sexual characteristics, such as the growth of body and facial hair, deepening of the voice, and increased sexual desire. Some less desirable effects are an increase in oiliness of the skin and acne.

Testosterone production in men tends to increase during childhood and reaches a maximum during the late teens or early twenties. It then decreases throughout the rest of a man's life. A sudden or extreme decrease in testosterone levels, caused by disease or injury to the hypothalamus, pituitary gland, or testes, can lead to a medical condition known as hypogonadism.

Treatments for hypogonadism in the form of testosterone injections, tablets, skin patches, and skin gels are available. Some men use these devices in an attempt to revive masculine traits that begin to decline normally as one grows older. Many individuals believe, like Brown-Séquard, that testosterone can be something of a "miracle drug" that will restore their lost youth. It has at times been recommended also as a treatment for a host of medical problems, including infertility, impotence, lack of sex drive, osteoporosis,

Words to Know

ANDROGENIC STEROID A hormone responsible for the development of male sexual characteristics.

DERIVATIVE A chemical whose structure is based on or related to another chemical.

HORMONE A chemical produced by living cells in a body that promotes the activity of other cells in the body.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

shortness of stature, anemia, and low appetite. Testosterone may or may not be helpful in treating any one of these conditions.

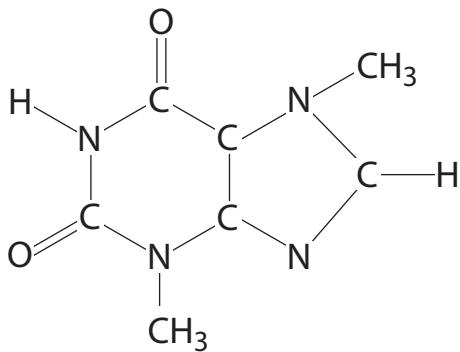
Since the 1950s, athletes have been using testosterone, its derivatives, and related compounds to improve their performance. The compound increases a person's bone and muscle mass, significantly improving his or her strength and endurance. Testosterone was first used on a large scale basis by athletes from the Soviet Union in the 1950s as part of that nation's efforts to become dominant in world sports.

One problem with using testosterone as a performance-enhancing drug is its undesirable side effects. It tends to increase the size of a man's prostate gland and decrease the size of his testicles. It may also produce wide mood swings that may include dangerously aggressive feelings.

Because of these side effects, researchers have developed compounds that produce the same effects as those obtained from natural testosterone, but without the compound's harmful side effects. Some of those compounds increase the natural production of testosterone in the body when they are ingested. Others are derivatives of testosterone, compounds with similar chemical structures, but minor changes that reduce side effects. These derivatives include compounds such as dihydrotestosterone, androstenedione (also known as andro), dehydroepiandrosterone (DHEA), clostebol, and nandrolone. Today most derivatives of testosterone and testosterone-producing compounds are banned by sports organizations, both because of their harmful side effects and because of the unfair advantages they provide athletes who use them.

FOR FURTHER INFORMATION

- Hellstrom, Wayne J. G. "Testosterone Replacement Therapy." *Digital Urology Journal*.
<http://www.duj.com/Article/Hellstrom2/Hellstrom2.html>
(accessed on November 15, 2005).
- "Material Safety Data Sheet: Testosterone." Paddock Laboratories, Inc.
<http://www.paddocklabs.com/forms/msds/testost.pdf>
(accessed on November 15, 2005).
- "Testosterone." International Programme on Chemical Safety.
<http://www.inchem.org/documents/pims/pharm/pim519.htm>
(accessed on November 15, 2005).
- "Testosterone Deficiency." Urology Channel.
<http://www.urologychannel.com/testosteronedeficiency/index.shtml> (accessed on November 15, 2005).



OTHER NAMES:

3,7-Dimethylxanthine;
3,7-dihydro-3,7-dimethyl-1H-purine-2,6-dione

FORMULA:

C₇H₈N₄O₂

ELEMENTS:

Carbon, hydrogen, nitrogen, oxygen

COMPOUND TYPE:

Alkaloid (organic base)

STATE:

Solid

MOLECULAR WEIGHT:

180.16 g/mol

MELTING POINT:

Begins to sublime at about 290°C (554°F)

BOILING POINT:

Not applicable

SOLUBILITY:

Slightly soluble in water and ethyl alcohol; virtually insoluble in benzene, chloroform, ether, and other organic solvents

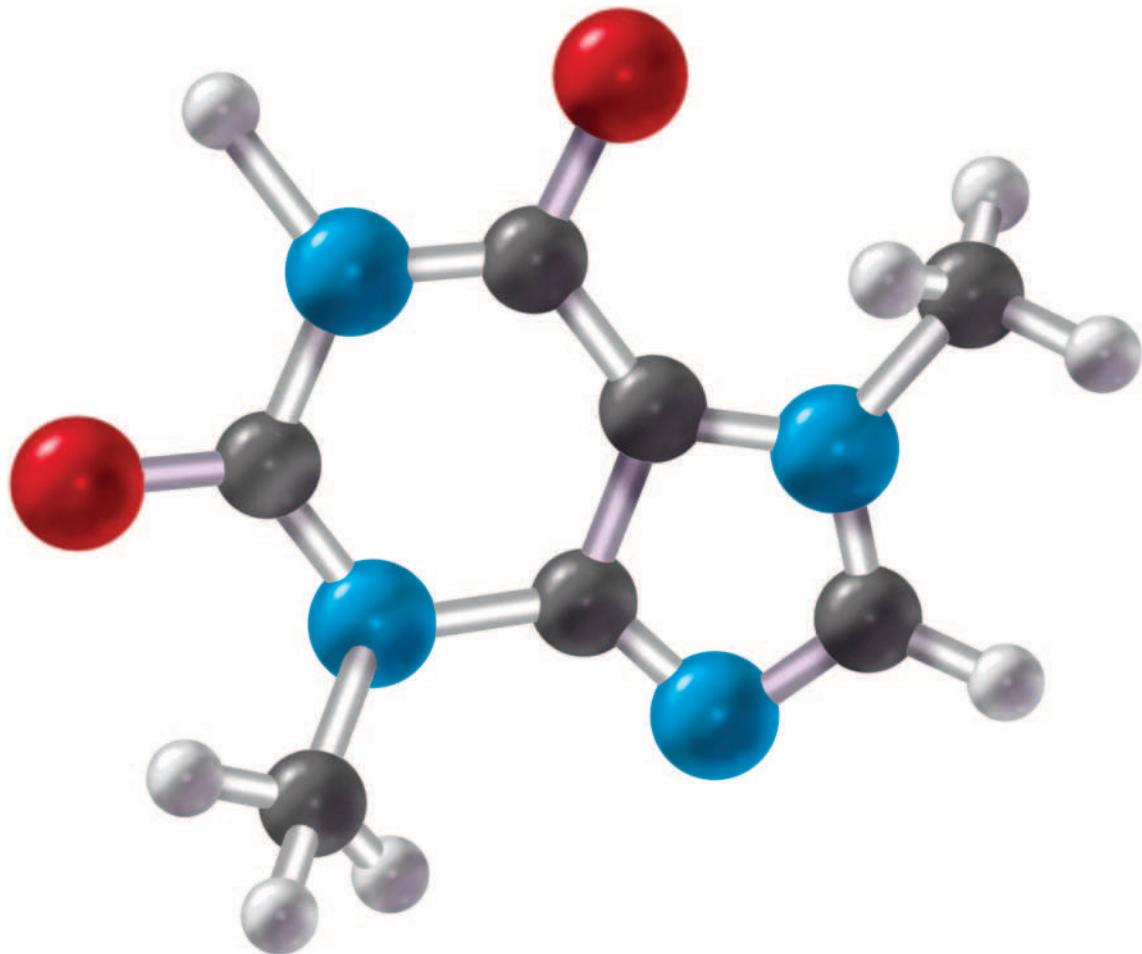
K
E
Y
F
A
C
T
S

OVERVIEW

Theobromine (thee-oh-BROH-meen) is a white crystalline solid that occurs naturally in cocoa beans, from which chocolate is obtained, and, in smaller amounts, in tea and cola nuts. Theobromine is structurally very similar to caffeine, which differs only in the presence of a methyl group (-CH₃) on one of the nitrogen atoms in the theobromine molecule. Both theobromine and caffeine belong to a family of organic compounds known as the methylxanthines. Theobromine's effects on the human body are similar to those of caffeine, but about ten times weaker. In addition, caffeine is metabolized more quickly, is addictive, and increases alertness and emotional stress. It may also have serious effects on the central nervous system and the kidneys. By contrast, theobromine produces feelings of well-being, is not addictive, has no effect on the central nervous system, and provides only gentle stimulation to the kidneys. Its effects on the body are much longer-lasting than are those of caffeine.

The amount of theobromine in cocoa beans varies widely, ranging from 10 to 40 milligrams of theobromine per gram

THEOBROMINE



Theobromine. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and blue atoms are nitrogen. Gray sticks indicate double bonds.

PUBLISHERS RESOURCE GROUP

of cocoa. The variation depends on a number of facts, including the type of bean, the location where it was grown, and the method of processing the bean. All chocolate products contain theobromine, but the amount varies depending on the type of chocolate. Dark chocolate contains significantly more of the compound than milk chocolate, and high quality chocolate tends to contain more theobromine than low quality chocolate. The characteristic bitter taste of dark chocolate is due to the theobromine present in it.

HOW IT IS MADE

Theobromine is usually obtained from the hulls of cocoa beans left over after the production of chocolate. The hulls

Interesting Facts

- The name *theobromine* is derived from the scientific name for the cacao tree, *Theobroma cacao*.
- Spanish armies that invaded South and Central America in the sixteenth century used chocolate as a source of energy.
- Theobromine is harmless to humans, but very toxic to dogs, horses, and other domestic animals.

are crushed and then treated with an absorbent, such as water or liquid carbon dioxide, which dissolves the theobromine. The water or carbon dioxide is then allowed to evaporate, permitting the crystallization of the pure compound.

COMMON USES AND POTENTIAL HAZARDS

Theobromine occurs in all chocolate products. The pure compound has relatively few uses, however, most of them medical. For example, it has been used as a diuretic—a compound that increases the rate at which liquids are eliminated from the body—and as a mild stimulant. It has also been used to treat hypertension (high blood pressure) because it is a vasodilator, a compound that causes blood vessels to relax

Words to Know

DIURETIC A compound that increases the rate at which liquids are eliminated from the body.

METHYLXANTHINE A compound that is derived from xanthine ($C_5H_4N_4O_2$), with methyl groups (CH_3) replacing one or more of the hydrogen atoms.

STIMULANT: A substance that increases the activity of a living organism or one of its parts.

SUBLIME To go from solid to gaseous form without passing through a liquid phase.

THEOBROMINE

and expand in size. Theobromine also appears to be effective as a cough suppressant, although the quantities needed to achieve useful effects are quite large.

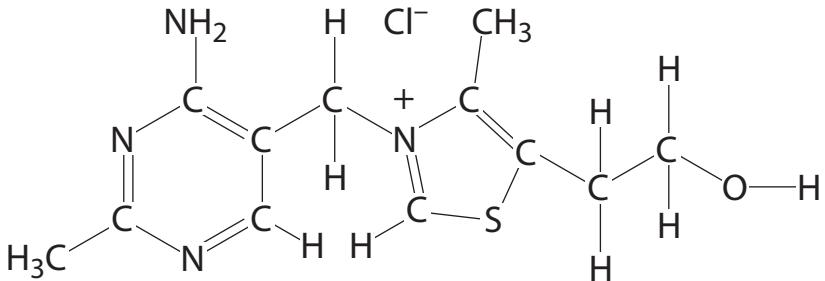
FOR FURTHER INFORMATION

O'Neil, John. "And It Doesn't Taste Bad, Either." *New York Times* (November 30, 2004): F9.

"Theobromine@3Dchem.com." Molecule of the Month.
<http://www.3dchem.com/molecules.asp?ID=155> (accessed on November 15, 2005).

"What Is Theobromine and What Is Its Effect on Human Beings?" International Cocoa Organization.
<http://www.icco.org/questions/theobromine.htm> (accessed on January 11, 2006).

See Also Caffeine; Carbon Dioxide



OTHER NAMES:

Vitamin B₁

FORMULA:
C₁₂H₁₇ClN₄OS

ELEMENTS:

Carbon, hydrogen, chlorine, nitrogen, oxygen, sulfur

COMPOUND TYPE:
Organic

STATE:
Solid

MOLECULAR WEIGHT:
300.82 g/mol

MELTING POINT:
Not applicable

BOILING POINT:
Not applicable

SOLUBILITY:
Soluble in water;
partially soluble in
ethyl alcohol

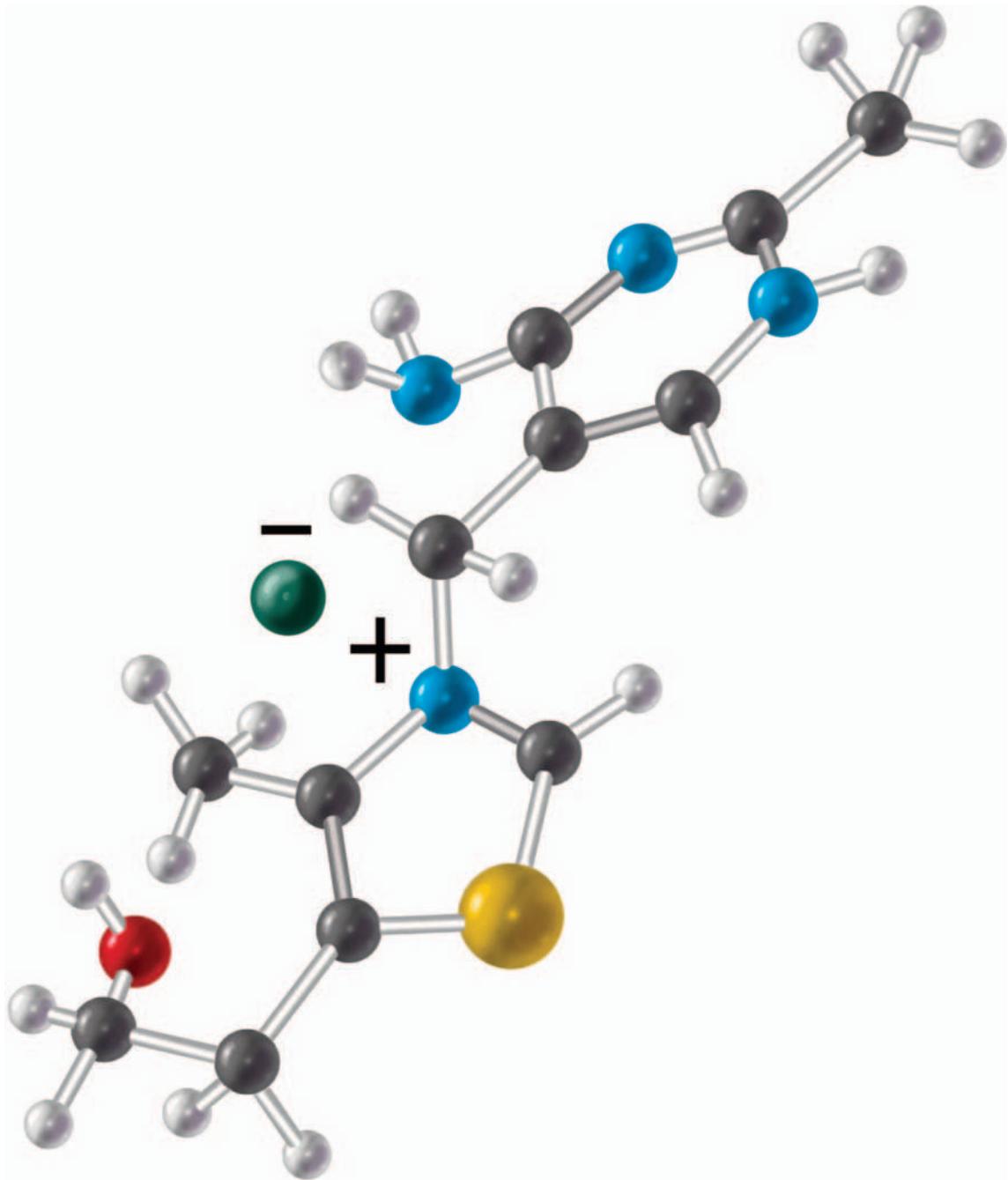
K
E
Y
F
A
C
T
S

OVERVIEW

Thiamine (THYE-uh-min) is the water-soluble vitamin, vitamin B₁. It is also known as 3-(4-Amino-2-methylpyrimidyl-5-methyl)-4-methyl-5,β-hydroxyethylthiazolium. It is usually made available as one of its salts, especially thiamine hydrochloride or thiamine mononitrate. Both salts are white crystalline solids with a bitter taste that are destroyed by alkaline solutions. An alkaline solution is one that contains a strong base.

Thiamine was discovered by Japanese scientist Suzuki Umetaro (1874–1943) in the early twentieth century. Umetaro was investigating a disease known as beriberi that had plagued humans for thousands of years. He found that the disease could be cured by feeding patients a diet that contained rice bran. He was also able to isolate a specific compound in rice bran that produced that effect, a compound he named aberic acid. Aberic acid later became known as thiamine.

The first written reports of beriberi date back 4,000 years. Chinese scholars described a disorder characterized by nausea,



Thiamine. Red atom is oxygen; white atoms are hydrogen; black atoms are carbon; blue atoms are nitrogen; and yellow atom is sulfur. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

vomiting, constipation, abdominal pain, weakness, wasting, paralysis, irritability, paranoia, depression, and other mental symptoms. In the 1800s, most scientists were convinced that bacteria were responsible for the disease. Over time, it became apparent that this theory was incorrect and that dietary factors

Interesting Facts

- The word beriberi comes from the Sinhalese expression “I can’t, I can’t.” Sinhalese is the language spoken by natives of Sri Lanka.
- As of 2006, Beriberi is rare in developed countries because of the ready availability of foods and

vitamin supplements that contain vitamin B₁. The disease became common in Cuba between 1989 and 1995, however. A trade embargo imposed by the United States reduced the availability of foods and vitamins for the Cuban people.

were responsible for beriberi. In the late nineteenth century, for example, Japanese naval scientist Kanehiro Takaki (1849–1920) noticed that sailors on long voyages developed beriberi if their diets consisted mainly of rice with hulls removed. He also observed that the sailors remained healthy if they included fish, vegetables, wheat, and milk along with rice in their diets.

In 1896, Dutch physician Christiaan Eijkman (1858–1930) found that animals developed beriberi only if they ate polished rice—rice from which hulls had been removed. He learned that the animals recovered if they ate rice with hulls still on it. His colleague Gerrit Grijns (1865–1944) believed that the rice hulls contained some substance that prevented beriberi. By 1910, Umetaro had found that substance, a-beric acid.

HOW IT IS MADE

Humans are unable to synthesize thiamine, so they must obtain the vitamin from other sources in which it occurs naturally. The best sources of thiamine are brewer’s yeast, whole grains, wheat germ, lean meats, organ meats (such as liver and kidney), fish, dried beans, soybeans, peas, nuts, green leafy vegetables, avocados, raisins, plums, and kelp. Thiamine is also made synthetically by a process that was first developed in the 1930s by American chemist Robert R.

Words to Know

COENZYME A chemical compound that works along with an enzyme to increase the rate at which chemical reactions take place.

SALT An ionic compound where the anion is derived from an acid.

Williams (1886-1965). Williams began his studies of anti-beriberi compounds in 1911 and worked for twenty-five years before discovering the correct chemical formula for the compound. Once he had that information, he was able to invent a system for producing the vitamin artificially. Today, virtually all of the vitamin B₁ sold commercially is prepared artificially.

COMMON USES AND POTENTIAL HAZARDS

Thiamine is a coenzyme needed for a number of essential biochemical reactions in the body. A coenzyme is a chemical compound that works along with an enzyme to increase the rate at which chemical reactions take place. Without enzymes and coenzymes in the body, many chemical reactions would take place so slowly that normal bodily functions could not continue. Thiamine is involved in chemical reactions by which blood is produced and circulated through the body, carbohydrates are metabolized, digestive enzymes are produced, and the nervous system is maintained.

Because it is soluble in water, thiamine is not stored in the body. A person must include the compound in his or her daily diet on a regular basis. Most people ingest adequate amounts of vitamin B₁ in their ordinary diets, and beriberi is very rare in developed countries of the world. It may occur, however, in alcoholics, pregnant women, and people who must undergo kidney dialysis. In all of these cases, a person does not receive adequate amounts of the vitamin for the body's needs. In the case of alcoholics, for example, alcohol replaces the calories they would be getting from food if they were not drinking so much. As a result, they do not get enough vitamin B₁ and other nutrients needed to stay healthy.

The early symptoms of beriberi include fatigue, irritability, poor sleep habits, memory loss, abdominal pain, loss of appetite, and constipation. As the disease progresses, it may produce damage to the peripheral nervous system that serves the arms, legs, feet, and hands, resulting in muscular atrophy (weakness and wasting of the muscles) and loss of sensation in the toes.

FOR FURTHER INFORMATION

Eades, Mary Dan. *The Doctor's Complete Guide to Vitamins and Minerals*. New York: Dell, 2000.

"Thiamine." Medline Plus.

<http://www.nlm.nih.gov/medlineplus/ency/article/002401.htm> (accessed on November 15, 2005).

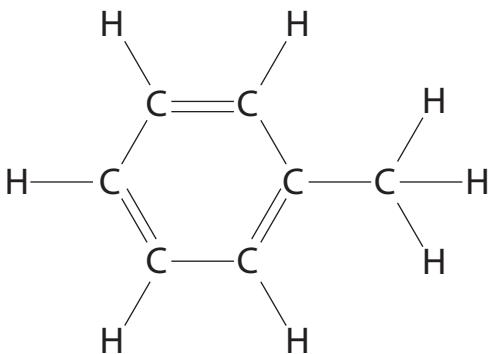
"Thiamine and Salts." International Programme on Chemical Safety.

<http://www.inchem.org/documents/pims/pharm/pimgo15.htm> (accessed on November 15, 2005).

"Vitamin B1." Welcome to Glactone.

<http://chemistry.gsu.edu/glactone/vitamins/b1/> (accessed on November 15, 2005).

See Also Ascorbic Acid; Cyanocobalamin; Pyridoxine; Riboflavin



OTHER NAMES:

Methylbenzene;
phenylmethane;
toluol

FORMULA:

$C_6H_5CH_3$

ELEMENTS:

Carbon, hydrogen

COMPOUND TYPE:

Aromatic hydrocarbon (organic)

STATE:

Liquid

MOLECULAR WEIGHT:

92.14 g/mol

MELTING POINT:

$-94.95^{\circ}C$ ($-138.9^{\circ}F$)

BOILING POINT:

$110.63^{\circ}C$ ($231.13^{\circ}F$)

SOLUBILITY:

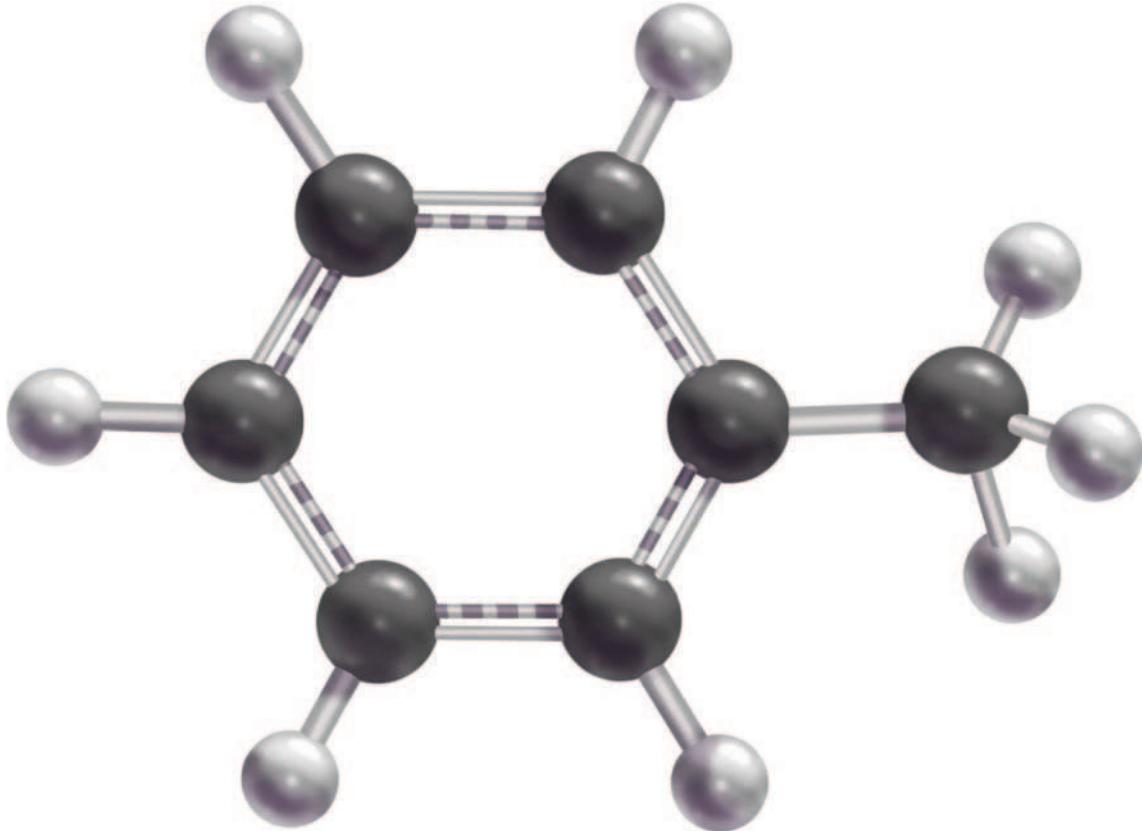
Insoluble in water;
miscible with ethyl
alcohol, ether,
acetone, and carbon
disulfide

K E Y F A C T S

Toluene

OVERVIEW

Toluene (TOL-yew-een) is a clear, colorless liquid with a benzene-like odor. It is an aromatic hydrocarbon, that is, a compound that contains carbon and hydrogen only, with the carbon atoms arranged in a ring. Aromatic compounds have structures based on that of benzene (C_6H_6). Toluene was discovered in 1838 by French chemist Pierre Joseph Pelletier (1788–1842). Pelletier found the compound in the gas emitted by the bark of the pine tree *Pinus maritima*. Pelletier named the substance *retinnaphte*, after pine resin. The compound was re-discovered a number of times in later years and given a variety of names, including heptacarbure quadihydrique, benzoene, and dracyl. Toluene's chemical nature was finally determined by German chemist August Wilhelm von Hofmann (1818–1892) and British chemist James Muspratt (1793–1886) who adopted the name *toluol* originally proposed by Swedish chemist Jons Jakob Berzelius (1779–1848).



Toluene. Black atoms are carbon; white atoms are hydrogen. Bonds in the benzene ring are represented by the double striped sticks. White sticks show single bonds.

PUBLISHERS RESOURCE GROUP

HOW IT IS MADE

Toluene occurs naturally in petroleum and is extracted during the catalytic reforming of the naphtha segment of petroleum. Catalytic reforming is a process by which hydrocarbons in petroleum are heated, often in contact with a catalyst, in order to change their molecular composition. The naptha segment of petroleum consists of a complex mixture of hydrocarbons with boiling points between about 50°C and 200°C (120°F and 400°F). Toluene is separated from other hydrocarbons produced during the reforming process by adding a solvent, such as sulfolane ($c\text{-CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{SO}_2$; the “c” at the beginning means that this is a ring compound) or tetraethylene glycol ($\text{HO}(\text{CH}_2\text{CH}_2\text{O})_3\text{CH}_2\text{CH}_2\text{OH}$). The toluene dissolves in one or the other of these solvents, while most other hydrocarbons with which it is mixed do not. Toluene is also obtained by a similar process starting with

Interesting Facts

- The name toluol (and toluene) originally comes from an aromatic substance obtained from the bark of the South American tree *Myroxylon balsamum*.
- Toluene was first synthesized artificially by German chemists Rudolf Fittig (1835–1910) and B. C. G. Tollens (1841–1918).
- Toluene is the recommended solvent for the removal of stains made by airplane cement.

coal tar oil, the liquid that remains when soft coal is heated in the absence of air to make coke.

COMMON USES AND POTENTIAL HAZARDS

Toluene generally ranks among the top thirty compounds produced in the United States annually. In 2002, American chemical plants produced 330 million liters (87 million gallons) of toluene for domestic use and export. About half that amount was used in the production of benzene and about 20 percent for the production of xylene ($C_6H_4(CH_3)_2$). Another 18 percent went to the manufacture of other organic chemicals, one of the most important of which is toluene diisocyanate ($CH_3C_6H_3(NCO)_2$) used primarily for the production of polyurethane foams and artificial rubber products. The other major use for toluene is as a solvent for paints and other coatings, gums, oils, natural and artificial rubber products, adhesives, and plastics. Some toluene is also used as a blending agent in aviation fuels and other gasoline mixtures with high-octane ratings.

One application of toluene that consumes relatively small amounts of the compound but that is very important is in the production of the explosive trinitrotoluene (TNT). TNT is a very desirable explosive because it detonates (explodes) only when shocked or heated to temperatures above 230°C (450°F), but then releases very large amounts of energy. For many decades, TNT was one of the world's most widely

Words to Know

AROMATIC COMPOUND A compound whose chemical structure is based on that of benzene (C_6H_6).

CATALYTIC REFORMING A process by which hydrocarbons in petroleum are heated, often in contact with a catalyst, in order to change their molecular composition.

MISCIBLE Able to be mixed; especially applies to the mixing of one liquid with another.

SOLVENT A substance that is able to dissolve one or more other substances.

TERATOGENIC Capable of causing an organism to develop incorrectly.

used explosives. Even today, the power of other explosives is measured by comparison with TNT. For example, the power of nuclear weapons is said to be 10 kilotons or 10 megatons, meaning that they have an explosive power comparable to 10,000 tons or 10 million tons, respectively, of TNT.

Toluene is a very hazardous chemical compound. It catches fire easily and, under proper conditions, is explosive. It is also an irritant to the skin, eyes, and respiratory tract. It may produce symptoms such as headache, dizziness, drowsiness, confusion, fatigue, and peculiar sensations of the skin, such as a “pins and needles” effect. In large quantities, toluene may cause unconsciousness, coma, and death. In contact with the skin or eyes, it may produce redness and pain. Prolonged exposure to toluene may cause serious damage to the liver and kidneys resulting in anemia, decreased blood cell count, and bone marrow hypoplasia (diminished ability to produce blood cells). Toluene is thought to be teratogenic, capable of damaging a fetus while still in the womb.

FOR FURTHER INFORMATION

Harte, John, et al. *Toxics A to Z*. Berkeley: University of California Press, 1991, 415–417.

Martin, Kevin A. “Toxicity, Toluene.” eMedicine.
<http://www.emedicine.com/emerg/topic594.htm> (accessed on December 29, 2005).

“Toluene.” International Chemical Safety Cards.

<http://www.cdc.gov/niosh/ipcsneng/neng0078.html> (accessed on December 29, 2005).

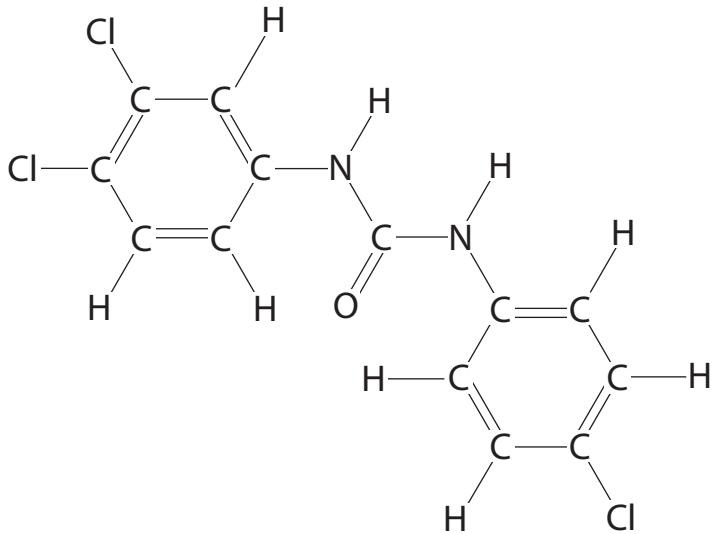
“Toluene.” J. T. Baker.

<http://www.jtbaker.com/msds/englishhtml/t3913.htm>
(accessed on December 29, 2005).

“ToxFAQs™ for Toluene.” Agency for Toxic Substances and Disease Registry.

<http://www.atsdr.cdc.gov/tfacts56.html> (accessed on December 29, 2005).

See Also Benzene



OTHER NAMES:

3,4,4'-Trichloro-carbanilide,
N-(4-chlorophenyl)-
N'-(3,4-dichlorophenyl)urea

FORMULA:

$C_6H_3Cl_2NH\text{-}CONHC_6H_4Cl$

ELEMENTS:

Carbon, hydrogen,
nitrogen, oxygen,
chlorine

COMPOUND TYPE:

Organic

STATE:

Solid

MOLECULAR WEIGHT:

315.58 g/mol

MELTING POINT:

256°C (493°F)

BOILING POINT:

Not applicable

SOLUBILITY:

Insoluble in water;
soluble in fats and
many organic
solvents

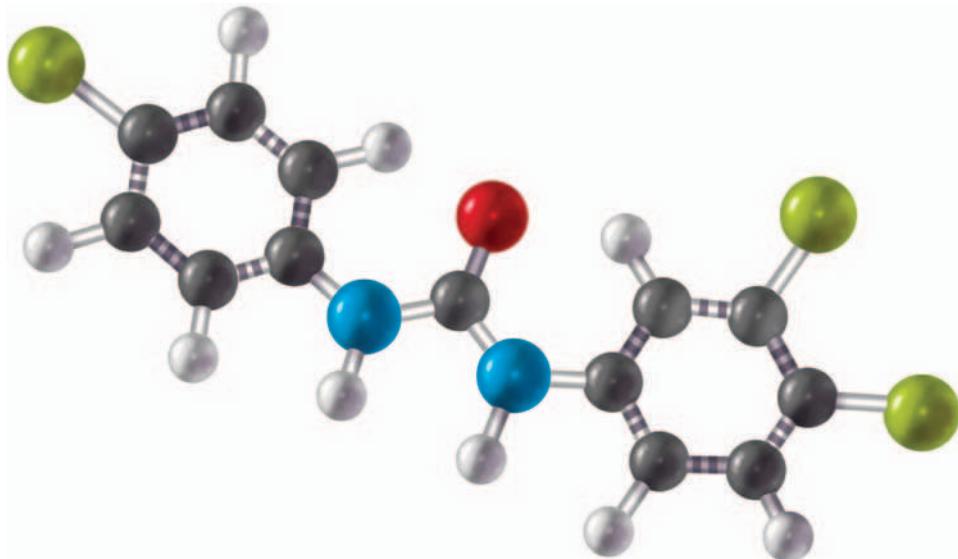
K
E
Y
F
A
C
T
S

Triclocarban

OVERVIEW

Triclocarban (TRY-klo-kar-ban) is a fine white powder with a slight odor. It is used almost exclusively as an antimicrobial agent—a substance that kills microorganisms such as bacteria and fungi. Triclocarban is available under a number of trade names, including Cusiter®, Cutisan®, Genoface®, Procutene®, Solubacter®, and TCC®. The compound is an ingredient in deodorant soaps, disinfectants, toothpastes, and other household products. It is also used in hospitals to cleanse and disinfect the skin of patients and medical staff. It remains on surfaces for relatively long periods of time, lengthening the disinfecting action, rather than breaking down fairly quickly like many other disinfectants. Triclocarban can be imbedded in plastics and textiles to create long-lasting antimicrobial surfaces, such as in toys and kitchenware. It is sometimes used in combination with triclosan, a similar antimicrobial agent.

Triclocarban works as a disinfectant because it disables an enzyme called enoyl-acyl-carrier-protein reductase, or



Triclocarban. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; green atoms are chlorine; and blue atoms are nitrogen. Gray sticks indicate double bonds; striped sticks indicate benzene rings.

PUBLISHERS RESOURCE GROUP

ENR, that many bacteria and fungi use to make cell membranes. If they cannot make cell membranes, these microorganisms die. Humans do not have this enzyme, so triclocarban is considered harmless to them. The use of triclocarban became more widespread after another popular antimicrobial, hexachlorophene, was banned by the U.S. Food and Drug Administration in the 1970s. Hexachlorophene use was discontinued after the antimicrobial was implicated in the death of infants who had been exposed to the product. In the first years of the twenty-first century, triclocarban and triclosan began to be used less, due to increased concerns about the long-term effects of these compounds on the human body and on the environment.

HOW IT IS MADE

Triclocarban is synthesized by one of two procedures. In one, 4-chlorophenyl isocyanate ($\text{ClC}_6\text{H}_4\text{NCO}$) is reacted with 3,4-dichloroaniline ($\text{C}_6\text{H}_3\text{NH}_2\text{Cl}_2$) to give triclocarban. The relationship of these two compounds to the structure of the final product ($\text{C}_6\text{H}_3\text{Cl}_2\text{NHCONHC}_6\text{H}_4\text{Cl}$) is obvious from their chemical structures. In the second method of preparation 3,4-dichlorophenyl isocyanate ($\text{Cl}_2\text{C}_6\text{H}_3\text{NCO}$) is reacted with 4-chloroaniline ($\text{ClC}_6\text{H}_4\text{NH}_2$) to give the desired product.

Interesting Facts

- Hunters sometimes wash their clothes in detergents that contain triclocarban to kill the bacteria that contribute to body odor. In this way, the animals on whom they prey, such as deer, are less likely to detect the smell of the hunter and flee the scene.
- No triclocarban is manufactured in the United States. It is imported from other countries, primarily China and India.

COMMON USES AND POTENTIAL HAZARDS

About three-fourths of all liquid soaps and nearly one-third of all bar soap made in the United States contains antimicrobial agents such as triclocarban and triclosan. Research has shown, however, that washing one's hand with plain soap and water provides as much protection from disease-causing organisms as do antimicrobial soaps. It is apparently the action of washing that removes most bacteria. Many public health experts say that antimicrobial soaps are not necessary in normal healthy households.

The growing concern is that the widespread use of antimicrobial soaps may cause harmful bacteria and fungi to become resistant to antimicrobial agents such as triclocarban and triclosan. That could happen if the genes in microorganisms that produce the enzyme ENR mutate, that is, change their chemical structure. Bacteria carrying the new gene might be more resistant to antimicrobial compounds.

Triclocarban has been shown to cause contact dermatitis, an allergic skin rash caused by contact with an irritating substance. This effect is of particular concern among infants and small children. The compound has been implicated in illnesses among newborns, and its use in maternity wards is discouraged. However, the compound is effective in treating other forms of skin rashes, such as atopic dermatitis, also known as eczema, because it kills the bacteria that cause the condition. Although the effects of ingesting triclocarban have not been thoroughly investigated, there is some question about the compound's possible carcinogenic effects.

Words to Know

ANTIMICROBIAL A substance that destroys microorganisms such as bacteria and fungi.

CARCINOGEN A chemical that causes cancer in humans or other animals.

ENZYME A complex protein that is produced by living cells and which promotes biochemical reactions in the body.

Triclocarban is considered by some authorities to be an environmental hazard, as well as a risk to human health. The compound is released into waterways when it is used for washing and bathing. Although wastewater treatment plants can remove up to 98 percent of the triclocarban in water, public health researchers have found the compound in 60 percent of the U.S. waterways studied. In one study, it was the fifth most common contaminant among 96 pollutants studied. The half life of triclocarban (the time it takes for half of the substance to disappear) is 1.5 years. That means that the compound will remain in water supplies for relatively long periods of time, certainly more than a few years. Although triclocarban is harmless to humans, it can be toxic to some types of aquatic life such as shellfish.

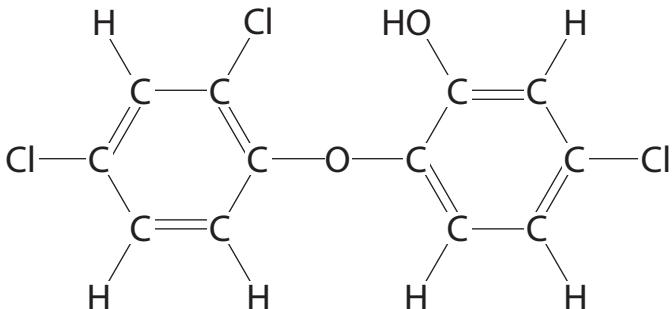
FOR FURTHER INFORMATION

“3,4,4’ Trichlorocarbanilide.” Hazardous Substances Data Bank.
<http://toxnet.nlm.nih.gov/cgi-bin/sis/search> (accessed on November 19, 2005).

“Anti bacterial Additive Triclocarban Widespread in U.S. Water ways.” Johns Hopkins Bloomberg School of Public Health, Public Health News Center.
http://www.jhsph.edu/PublicHealthNews/Press_Releases/2005/Halden_triclocarban_triclosan.html (posted on January 21, 2005; accessed January 6, 2006).

Senese, Fred. “What Are Triclocarban and Triclosan (Ingredients in Some Antiseptic Soaps)?” General Chemistry Online!
<http://antoine.frostburg.edu/chem/senese/101/consumer/faq/triclosan.shtml> (accessed on November 19, 2005).

See Also Triclosan



OTHER NAMES:

See Overview

FORMULA:

$\text{C}_6\text{H}_3\text{ClOH-O-C}_6\text{H}_3\text{Cl}_2$

ELEMENTS:

Carbon, hydrogen, oxygen, chlorine

COMPOUND TYPE:

Organic

STATE:

Solid

MOLECULAR WEIGHT:

289.54 g/mol

MELTING POINT:

55°C to 57°C (131°F to 135°F)

BOILING POINT:

Not applicable

SOLUBILITY:

Slightly soluble in water; soluble in many organic solvents

KEY FACTS

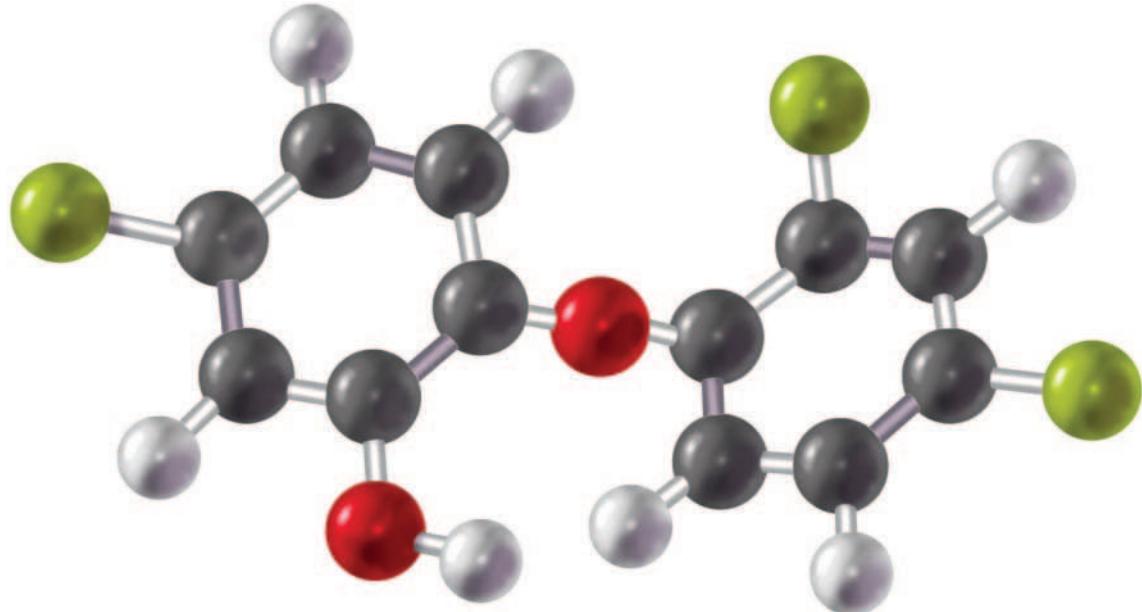
OVERVIEW

Triclosan (TRY-klo-san) is a white crystalline powder with antimicrobial properties that make it a useful ingredient in soaps, cosmetics, acne medications, deodorants, foot sprays and foot powders, toothpastes, and mouthwashes. It acts as an antimicrobial by inhibiting the action of an enzyme called enoyl-acyl carrier-protein reductase (ENR) that bacteria and fungi need to survive. The enzyme is used in the synthesis of fatty acids from which cell membranes are constructed. Having lost the ability to manufacture cell walls, bacteria and fungi die. The ENR enzyme is not present in humans, so triclosan has no effect on the human body.

Triclosan is also known by the following names: 5-chloro-2-(2,4-dichlorophenoxy)phenol; trichloro-2'-hydroxydiphenyl-ether; and 2,4,4'-trichloro-2'-hydroxydiphenyl ether.

HOW IT IS MADE

Two procedures are available for the synthesis of triclosan. In one, 1,4-dichloro-2-nitrobenzene ($\text{C}_6\text{H}_3\text{Cl}_2\text{NO}_3$) is



Triclosan. Red atoms are oxygen; white atoms are hydrogen; black atoms are carbon; and green atoms are chlorine. Dark gray sticks are double bonds; other sticks are single bonds. PUBLISHERS RESOURCE GROUP

reacted with 2,4-dichlorophenol (C_6H_3ClOH) to obtain triclosan. In the second procedure, 2,4-dichlorobenzene ($C_6H_4Cl_2$), acetyl chloride (CH_3OCl), and 2,4-dichlorophenol (C_6H_3ClOH) are reacted with each other to obtain the product.

COMMON USES AND POTENTIAL HAZARDS

Triclosan has been used in soaps and other cleaning products since the 1960s. Hospitals found the compound to be very effective in controlling bacteria on hands and surfaces. Medical staff used it to stop outbreaks of infections caused by common bacteria such as *Staphylococcus aureus*. The compound is still used widely in hospitals for this purpose. In the 1990s, triclosan began to appear in a much wider array of products including antibacterial soaps, body washes, toothpastes, and dishwashing liquids. By the early 2000s, it was also being used in toothbrushes, plastic toys, and socks. The addition of triclosan to tooth care products is based on evidence that the compound may inhibit the growth of bacteria that cause dental caries (cavities). The compound has also proved to be effective in killing the

Interesting Facts

- Soaking socks in a solution that contains triclosan can prevent food odor because the compound kills the

bacteria that grow on feet and are responsible for the odor.

parasites that cause malaria, one of the most serious diseases in the world.

In spite of its wide use, a number of questions have been raised about the inclusion of triclosan in household and commercial products. One objection is based on the concern that widespread use of the compound may lead to the development of bacteria resistant to antibiotics. Questions have also been raised about byproducts of the reactions by which triclosan is produced and by which it is degraded in the soil. Among those byproducts are a family of organic compounds known as the dioxins, among the most potent toxins known to humans. In addition, studies suggest that cleaning products that contain triclosan are not any more effective in killing germs than is the ordinary procedure of washing one's hands with soap and water. As a result, the Federal Trade Commission has ordered some companies to stop claiming that their product kills germs and reduces disease. The U.S. Environmental Protection Agency has listed triclosan as a possible human carcinogen.

Words to Know

CARCINOGEN A chemical that causes cancer in humans or other animals.

from simple beginning chemicals, or reactants.

SYNTHESIS A chemical reaction in which some desired product is made

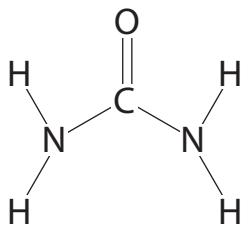
FOR FURTHER INFORMATION

Glaser, Aviva. "The Ubiquitous Triclosan." Beyond Pesticides.
<http://www.beyondpesticides.org/pesticides/factsheets/Triclosan%20cited.pdf> (accessed on November 19, 2005).

"Material Safety Data Sheet." Jeen International Corporation.
<http://www.jeen.com/cartexe/pdfs/msds%20JEECHEM%20TRICLOSAN.pdf> (accessed on November 19, 2005).

Mirsky, Steve. "Home, Bacteria Ridden Home: Could Antibacterial Soaps Lead to Resistant Strains?" *Scientific American* (July 19, 1997). Also available online at
<http://sciam.com/article.cfm?articleID=0009078D A418 1C76 9B81809EC588EF21> (accessed on November 19, 2005).

See Also Triclocarban

**OTHER NAMES:**

Carbamide;
carbonyldiamide

FORMULA:

$(\text{NH}_2)_2\text{CO}$

ELEMENTS:

Carbon, nitrogen,
hydrogen, oxygen

COMPOUND TYPE:

Organic

STATE:

Solid

MOLECULAR WEIGHT:

60.06 g/mol

MELTING POINT:

133.3°C (271.9°F)

BOILING POINT:

Not applicable;
decomposes above its
melting point

SOLUBILITY:

Soluble in water,
ethyl alcohol, and
benzene; slightly
soluble in ether

KEY FACTS

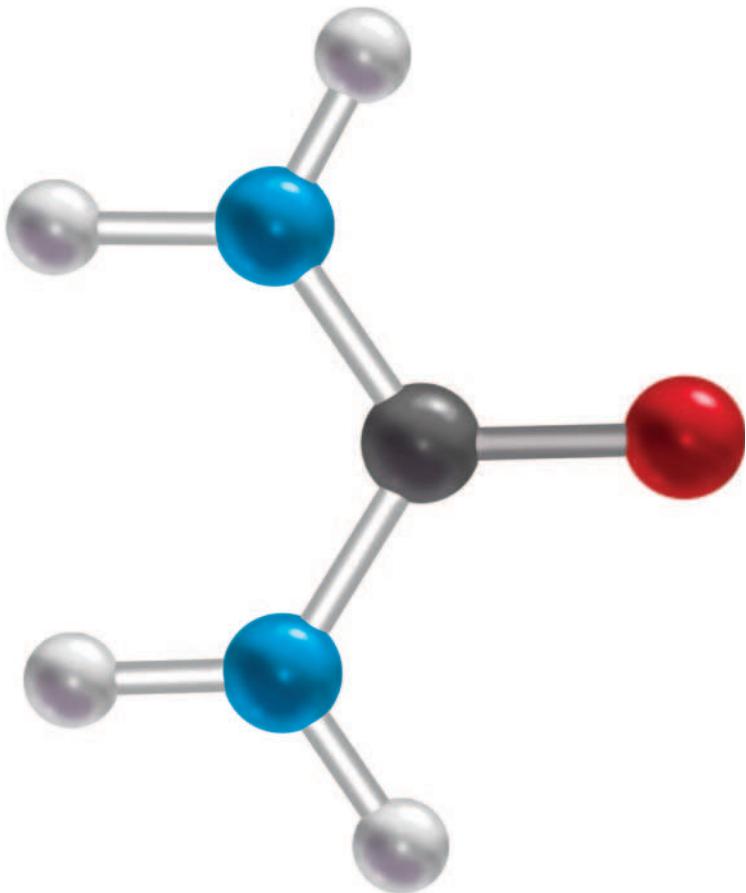
Urea

OVERVIEW

Urea (yoo-REE-uh) is a white crystalline solid or powder with almost no odor and a salty taste. It is a product of the decomposition of proteins in the bodies of terrestrial animals. Urea is produced in the liver and transferred to the kidneys, from which it is excreted in urine. The compound was first identified as a component of urine by French chemist Hilaire Marin Rouelle (1718–1799) in 1773. It was first synthesized accidentally in 1828 by German chemist Friedrich Wöhler (1800–1882). The synthesis of urea was one of the most important historical events in the history of chemistry. It was the first time that a scientist had synthesized an organic compound. Prior to Wöhler's discovery, scientists believed that organic compounds could be made only by the intervention of some supernatural force. Wöhler's discovery showed that organic compounds were subject to the same set of natural laws as were inorganic compounds (compounds for non-living substances). For this reason, Wöhler is often called the Father of Organic Chemistry.

Urea. Red atom is oxygen; white atoms are hydrogen; black atom is carbon; and blue atoms are nitrogen. Gray sticks indicate double bonds.

PUBLISHERS RESOURCE GROUP



HOW IT IS MADE

The formation of urea is the evolutionary solution to the problem of what to do with poisonous nitrogen compounds that formed when proteins decompose in the body. Proteins are large, complex compounds that contain relatively large amounts of nitrogen. When they decompose, that nitrogen is converted to ammonia (NH_3), a substance that is toxic to animals. If animals are to survive the decomposition of proteins (as happens whenever foods are metabolized), some method must be found to avoid the buildup of ammonia in the body.

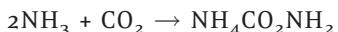
That method involves a series of seven chemical reactions called the urea cycle by which nitrogen from proteins

Interesting Facts

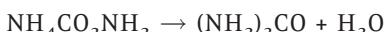
- Different species of animals have evolved different methods of eliminating ammonia from their bodies. For example, fish excrete the ammonia produced by the decomposition of proteins directly into the watery environment in which they live. Birds, who consume less water by gram of weight than do most other animals, convert ammonia to uric acid ($C_5H_4N_4O_3$), a white crystalline solid that is even less toxic than urea.
- Urea is transported from the liver to the kidneys in the bloodstream. By the time it leaves the body in urine, its concentration is sixty to seventy times its concentration in the bloodstream.

is converted into urea. Although high concentrations of urea do pose a risk to animal bodies, the urea formed in these reactions is normally excreted fast enough to avoid health problems for an animal.

Urea is produced commercially by the direct synthesis of liquid ammonia (NH_3) and liquid carbon dioxide (CO_2). The product of this reaction is ammonium carbamate ($NH_4CO_2NH_2$):



Ammonia and carbon dioxide do not react with each other under normal conditions of temperature and pressure. If the pressure is raised to 100 to 200 atmospheres (1750 to 3000 pounds per square inch) and the temperature is raised to about $200^\circ C$ ($400^\circ C$), however, the reaction proceeds efficiently with the formation of ammonium carbamate. When the pressure is then reduced to about 5 atmosphere (80 pounds per square inch), the ammonium carbamate decomposes to form urea and water:



Words to Know

METABOLISM All of the chemical reactions that occur in cells by which fats, carbohydrates, and other compounds are broken down to produce energy and the compounds needed to build new cells and tissues.

SYNTHESIS A chemical reaction in which some desired chemical product is made from simple beginning chemicals, or reactants.

COMMON USES AND POTENTIAL HAZARDS

Urea is the sixteenth most important chemical in the United States, based on the amount produced annually. In 2004, the chemical industry produced 5.755 million metric tons (6.344 million short tons) of urea. Almost 90 percent of that output was used in the manufacture of fertilizers. An additional 5 percent went to the production of animal feeds. In both fertilizers and animal feeds, urea and the compounds from which it is made provide the nitrogen needed by growing plants and animals for their good health and survival. The other major use of urea is in the manufacture of various types of plastics, especially urea-formaldehyde resins and melamine.

Urea is also used:

- In the production of personal care products, such as hair conditioners, body lotions, and dental products;
- In certain pharmaceutical and medical products, such as creams to treat wounds and damaged skin;
- As a stabilizer in explosives, a compound that places limits on the rate at which an explosion proceeds;
- In the manufacture of adhesives;
- For the flame-proofing of fabrics;
- For the separation of products produced during the refining of petroleum;
- In the production of sulfamic acid (HOSO_2NH_2), an important raw material in many chemical processes;
- As a coating for paper products; and
- In the production of deicing agents.

FOR FURTHER INFORMATION

National Urea Cycle Disorders Foundation.

<http://www.nucdf.org/> (accessed on November 19, 2005).

Ophardt, Charles E. "Urea Cycle." Virtual Chembook.

<http://www.elmhurst.edu/~chm/vchembook/633ureacycle.html> (accessed on November 19, 2005).

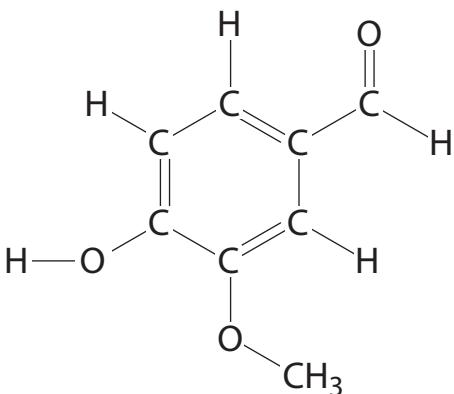
"Urea." International Programme on Chemical Safety.

<http://www.inchem.org/documents/icsc/icsc/eics0595.htm> (accessed on November 19, 2005).

"Urea." Third Millennium Online.

<http://www.3rd1000.com/urea/urea.htm> (accessed on November 19, 2005).

See Also Ammonia



OTHER NAMES:

4-hydroxy-3-methoxybenzaldehyde;
3-methoxy-4-hydroxybenzaldehyde; vanillic aldehyde

FORMULA:

(CH₃O)(OH)C₆H₃CHO

ELEMENTS:

Carbon, hydrogen, oxygen

COMPOUND TYPE:

Ether (organic)

STATE:

Solid

MOLECULAR WEIGHT:

152.15 g/mol

MELTING POINT:

81.5°C (179°F)

BOILING POINT:

285°C (545°F)

SOLUBILITY:

Slightly soluble in water; soluble in glycerol, ethyl alcohol, ether, and acetone

KEY FACTS

OVERVIEW

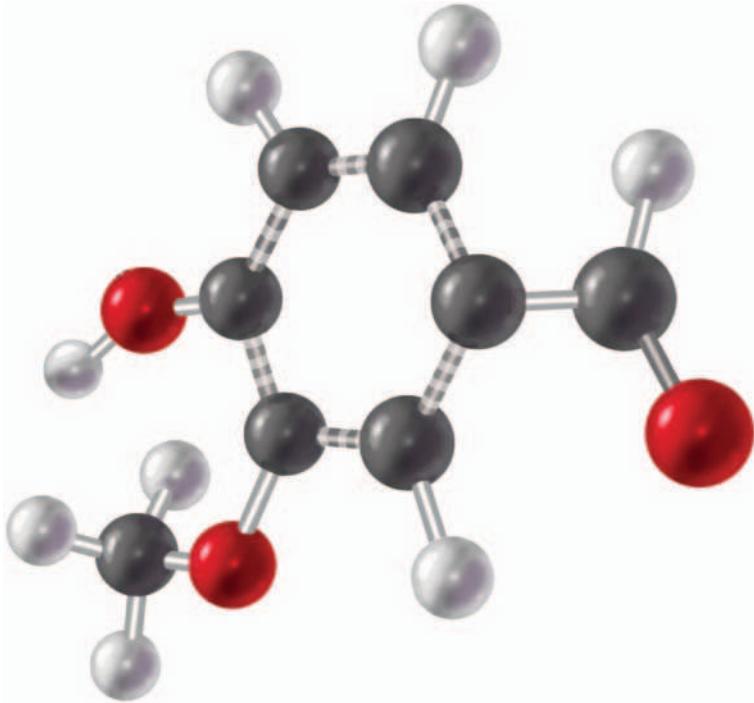
Vanillin is a white crystalline solid with a pleasant, sweet aroma, and a characteristic vanilla-like flavor. Chemically, it is the methyl ether of 4-hydroxybenzoic acid, a ring compound that contains the carboxyl (-COOH) group and the hydroxyl (-OH) group. Vanillin is the substance responsible for the familiar taste of vanilla, which has been used as a food additive and spice for hundreds of years. Vanilla was probably first used as a flavoring by the inhabitants of South and Central America before the arrival of Europeans in the sixteenth century. Spanish explorers brought the spice back to Europe, where it soon became very popular as a food additive and for the flavoring of foods. Since that time, vanilla has become one of the world's most popular spices.

HOW IT IS MADE

Vanilla is obtained naturally from the seed pod of the tropical orchid *Vanilla planifolia* by a lengthy and expensive

Vanillin. Red atoms are oxygen; white atoms are hydrogen; and black atoms are carbon. Gray stick indicates double bond; striped sticks indicate benzene ring.

PUBLISHERS RESOURCE GROUP



process. The pods are picked before they ripen and then cured until they are dark brown. The curing process involves soaking the pods in hot water, sun-drying them, and allowing them to “sweat” in straw. The cured pods are then soaked in alcohol to produce a product known as pure vanilla extract. The primary constituent in pure vanilla extract is vanillin, which gives the product its flavor. The process of extracting pure vanilla from seed pods may take as long as nine months.

Some people prefer a vanilla product that contains no, or almost no, alcohol. If alcohol is removed, almost pure vanilla is left behind, leaving a product known as natural vanilla flavoring.

Vanilla can also be extracted from plants other than *Vanilla planifolia*, such as potato peels and pine tree sap. The most economical source of the product, however, is waste material left over from the wood pulp industry. That waste material consists primarily of lignin, a complex natural polymer that, along with cellulose, is the primary component of wood. The wastes from wood pulping can be treated to break down and separate the lignin. This leaves behind a complex

mixture, a major component of which is vanilla. This vanilla is called lignin vanilla and has many of the same physical properties as natural vanilla. Since it is so much less expensive to make, it has become one of the major forms of vanilla used by consumers. Lignin vanilla is known commercially as artificial vanilla flavoring.

The two forms of vanilla described earlier—natural vanilla and lignin vanilla—are mixtures in which the compound vanillin is a major component. In both mixtures, other components are present in lesser amounts. These components may add somewhat different flavors and aromas, modifying the pure taste and smell of vanillin. Artificial methods for the production of pure vanillin have been available since the late 1890s. The most popular of those methods begins with eugenol ($(C_3H_5)C_6H_3(OH)OCH_3$) or isoeugenol ($(CH_3CHCH)C_6H_3(OH)OCH_3$). Either of these compounds is then treated with acetic anhydride ($(CH_3CO)_2O$) to obtain vanillin acetate, which is then converted to pure vanillin. The product of this reaction, unlike natural vanilla and lignin vanilla, is a pure compound, 4-hydroxy-3-methoxybenzaldehyde, pure vanillin. This method was the primary method for making artificial vanillin for more than 50 years. It has since been replaced by an alternative method of preparation, the Reimer-Tiemann reaction. This method for making artificial vanillin begins with catechol ($C_6H_4(OH)_2$) or guaiacol ($CH_3OC_6H_4OH$).

The Remier-Tiemann reaction is also used to produce another form of vanillin called ethyl vanillin. Ethyl vanillin is the ethyl ether of 4-hydroxybenzoic acid, 4-hydroxy-3-ethoxybenzaldehyde ($(CH_3CH_2O)(OH)C_6H_3CHO$). It is a close chemical relative of natural vanillin in which the methyl ($-CH_3$) group of natural vanillin is replaced by an ethyl ($-CH_2CH_3$) group. Ethyl vanillin is also known as artificial vanilla or synthetic vanilla. Its flavor is about three times as strong as that of methyl vanillin and is used to fortify or replace natural vanillin and lignin vanillin.

COMMON USES AND POTENTIAL HAZARDS

All forms of vanillin are used as a flavoring agent and sweetener in many types of foods, including candies, dessert products, ice creams, puddings, yogurts, diet shakes, and soft drinks. It is also added to some wines, alcoholic liquors,

Interesting Facts

- The tropical orchid *Vanilla planifolia* is pollinated naturally by the tiny Melipone bee, native to Mexico. In areas where the bee does not live, the orchid must be pollinated artificially by humans.
- Less than 1 percent of the vanillin produced annually comes from vanilla beans. The remaining 99 percent comes from lignin or is produced by synthetic means.
- Small amounts of vanillin are present in the wood used to make wine casks and adds to the flavor of wine.
- The Coca-Cola Company is believed to be the world's largest buyer of pure vanillin. Since the company does not reveal the recipe for its products, that assumption cannot be confirmed.
- The word *vanilla* comes from the Spanish word *vainilla*, meaning "little sheath," which refers to the shape of the vanilla orchid.

toothpastes, and cigarettes. The vanillins have also been shown to stimulate one's appetite, so they have been used to treat appetite loss. They are also added to cattle feed to enhance weight gain.

However, less than half the vanillin produced is used in food products. Vanillin's rich fragrance makes the compound useful also as an additive in perfumes, air fresheners, soaps, shampoos, candles, creams, lotions, colognes, and ointments. The compound is also used as a raw material in the manufacture of a variety of drugs, particularly the compound known as L-dopa, used to treat Parkinson's disease.

Vanillin is considered safe for human consumption, although it can be toxic in very large quantities. Known reactions include respiratory irritation, including coughing and shortness of breath, and gastrointestinal tract irritation. Contact with the skin or eyes can also cause irritation, redness, and pain. These symptoms are virtually unknown except for individuals who work directly with the pure compounds.

FOR FURTHER INFORMATION

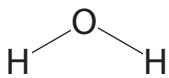
"All about Vanilla Extracts and Flavors." The Vanilla.COMpany.
<http://www.vanilla.com/html/facts extracts.html> (accessed on November 19, 2005).

"Aroma Chemicals from Petrochemical Feedstocks." National Economic and Development Council of South Africa.
<http://www.nedlac.org.za/research/fridge/aroma/part3/benchmarking.pdf> (accessed on November 19, 2005).

"Food Guide Question." Gloucester Muslim Welfare Association.
<http://www.gmwa.org.uk/foodguide2/viewquestion.php?foodqid=69&catID=1&compID=1> (accessed on November 19, 2005).

Rain, Patricia. *Vanilla: A Cultural History of the World's Most Popular Flavor and Fragrance*. Edited by Jeremy P. Tarcher. New York: Penguin Group USA, 2004.

"Vanillin." Greener Industry.
http://www.uyseg.org/greener_industry/pages/vanillin/1Vanillin_AP.htm (accessed on November 19, 2005).



OTHER NAMES:
Hydrogen oxide;
dihydrogen oxide

FORMULA:
 H_2O or HOH

ELEMENTS:
Hydrogen, oxygen

COMPOUND TYPE:
Inorganic

STATE:
Liquid

MOLECULAR WEIGHT:
18.02 g/mol

MELTING POINT:
0°C (32°F)

BOILING POINT:
100°C (212°F)

SOLUBILITY:
Soluble in ethyl
alcohol, methyl
alcohol, and acetone

K E Y F A C T S

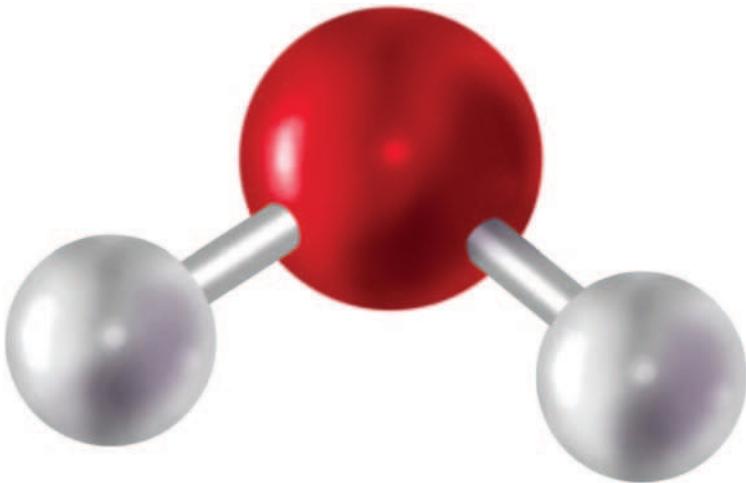
OVERVIEW

Water is a colorless, odorless, tasteless liquid that also occurs commonly in the solid state (as ice) and in the gaseous state (as steam or water vapor). It is a very stable compound that undergoes a number of important reactions. It reacts with some metals to form elemental hydrogen and an inorganic base. With some metals, such as sodium and potassium, the reaction is quite violent. With other metals, such as iron, the reaction occurs only very slowly. Water is also a weak electrolyte, ionizing to form hydrogen ions (H^+) and hydroxide ions (OH^-). The hydrogen ions occur in solution as hydronium ions (H_3O^+).

Water is also a strong dipole. A dipole is a molecule in which electrical charges are sufficiently separated from each other that one part of the molecule is negatively charged and another part, positively charged. Water's dipole character is responsible for many of its special characteristics. The positive end of one water molecule is attracted to the negative end of a second water molecule, resulting in the formation of a weak ("hydrogen") bond between the two molecules.

Water. Red atom is oxygen and white atoms are hydrogen.

PUBLISHERS RESOURCE GROUP



For example, water has a very high boiling point for a substance with relatively small molecules. The high boiling point is a result of the fact that heat added to water must first be used to break hydrogen bonds between water molecules before providing enough energy to vaporize the molecules. Similarly, the phenomenon known as surface tension is caused by hydrogen bonding. Surface tension is the tendency of a liquid to act as if it is covered with a thin film. Some insects are able to walk on water because its surface tension is so great. The surface tension is caused by the attractive forces between adjacent water molecules.

Water is also an excellent solvent. A solvent is a substance capable of dissolving other substances. Chemists sometimes refer to water as “the universal solvent” because it is able to dissolve so many other substances. That statement is an exaggeration, but does reflect the compound’s ability to dissolve more substances than probably any other single compound. Water’s ability to dissolve other substances is at least partly a result of its strong dipole character. The positive or negative end of a water molecule attaches itself to the negative or positive end of the substance to be dissolved. The force of attraction exerted by the water molecule is sufficient to tear apart the particles of which the second substance is composed causing them to dissolve in the water.

Interesting Facts

- When water freezes, the attractive forces between molecules force them into a regular crystalline pattern that occupies more space than do the same water molecules in the liquid state. In other words, water expands when it freezes and ice is less dense than liquid water.
- One consequence of that fact is that lakes freeze from the top down. As ice forms, it floats to the surface of the liquid. Aquatic organisms in the lake are able to live through the winter because of the layer of ice on top of the lake.
- Adding to the protection added by the layer of ice is the fact that ice is one of the best thermal (heat) insulators known. That is, heat flows through ice more slowly than through almost any other substance. Any heat left in a freezing lake will be conserved within the water since it escapes so slowly through the ice layer.

HOW IT IS MADE

Water can be made by a variety of chemical reactions, including:

- The oxidation of hydrogen: $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$;
- The reaction between an acid and a base, as, for example: $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$;
- The combustion of an organic material, as, for example: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$.

Because water occurs so abundantly, none of these reactions is required for the commercial production of the compound. Water makes up about 70 percent of the Earth's surface in the oceans, lakes, rivers, ponds, glaciers, ice caps, and other reservoirs. The problem is that only a very small fraction of that water—about 3 percent—is fresh water. The remaining 97 percent is salt water. And even the 3 percent of fresh water in lakes, rivers, and other resources is impure, in the sense that it contains other substances dissolved and suspended in it.

Thus, the primary concern in obtaining adequate supplies of pure water for household, personal, commercial, industrial, or other uses is the purification of water, not its synthesis. Purification of water is achieved by a number of processes, including chlorination, filtration, distillation, or purification by some type of ion exchange mechanism.

COMMON USES AND POTENTIAL HAZARDS

One of the most important uses of water is the survival of life on Earth. All plants and animals contain a high proportion of water. That proportion ranges from as high as 97 percent in many fruits and vegetables to a low of about 20 percent in some "dry" foods like breads and cereals. The human body itself comprises anywhere from 50 to 70 percent of a person's body weight. That water plays a number of roles, such as providing a solvent by which nutrients are circulated throughout the body and making possible all kinds of chemical reactions that occur in aqueous solutions, but do not occur in the dry state.

Although easy to ignore, water plays an almost unlimited number of roles in industrial, chemical, commercial, and other operations. Among these applications of water are:

- As a coolant in electricity generating plants, petroleum refineries, chemical plants, and many kinds of industrial factories;
- For irrigation;
- For cleaning, washing, and scouring raw materials and finished products;
- As a source of hydrogen and oxygen for many industrial and chemical operations;
- As a solvent for many types of industrial and chemical reactions and for the extraction of compounds from mixtures;
- In the manufacture of many kinds of foods and beverages, such as beer, wine, and soft drinks;
- As a medium for suspending materials in industrial processes, such as the manufacture of paper;
- In the processing of textiles;
- For the generation of steam to power industrial, household, and chemical processes;

Words to Know

AQUEOUS SOLUTION A solution that consists of some material dissolved in water.

ELECTROLYTE A substance which, when dissolved in water, will conduct an electric current.

HYDROLYSIS The process by which a compound reacts with water to form two new compounds.

- For the hydration of lime;
- As a coolant in nuclear reactors;
- In the transport of industrial and chemical raw materials and products;
- For the removal of barks from logs in the timber industry;
- To make possible hydrolysis reactions in chemical and industrial operations;
- In the manufacture of Portland cement; and
- To dilute solutions that are too concentrated for some given industrial process.

FOR FURTHER INFORMATION

“Iowa Project WET.” Iowa Academy of Science.

<http://www.uni.edu/~iowawet/iowawet.html> (accessed on November 19, 2005).

Strange, Veronica. *The Meaning of Water*. Oxford, UK: Berg Publishers, 2004.

“Water Properties.” London South Bank University.

<http://www.lsbu.ac.uk/water/data.html> (accessed on November 19, 2005).

“Water Science for Schools.” U.S. Geological Survey.

<http://ga.water.usgs.gov/edu/> (accessed on November 19, 2005).

OTHER NAMES:	Chinese white; philosopher's wool; zinc white; flowers of zinc
FORMULA:	ZnO
ELEMENTS:	Zinc, oxygen
COMPOUND TYPE:	Metallic oxide (inorganic)
STATE:	Solid
MOLECULAR WEIGHT:	81.39 g/mol
MELTING POINT:	1974°C (3585°F)
BOILING POINT:	Not applicable
SOLUBILITY:	Insoluble in water; soluble in dilute acids

KEY FACTS

Zinc Oxide

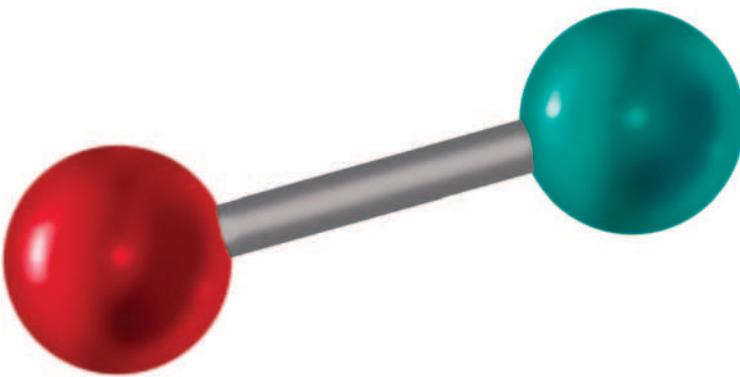
OVERVIEW

Zinc oxide (ZINC OX-side) is a white to gray to yellowish powder with no odor but a bitter taste. It occurs in nature as the mineral zincite. Zinc oxide is perhaps best known as an ingredient in cosmetics, personal care products, medicines, and sunscreens. In terms of volume, however, a number of industrial applications are of greater importance.

HOW IT IS MADE

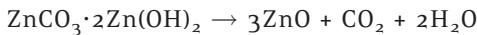
Three commercial methods of preparation are available for zinc oxide. Two of these methods are known as thermal procedures because they involve heating either zinc metal or an ore of zinc to obtain zinc vapor. The zinc vapor is then reacted with oxygen to produce zinc oxide. The formula for this reaction is $2\text{Zn} + \text{O}_2 \rightarrow 2\text{ZnO}$

The third method is called a “wet” procedure because it takes place in aqueous solution. In this process, zinc compounds are leached out of ores and converted to basic zinc



Zinc oxide. Red atoms are oxygen and turquoise atoms are zinc. Gray sticks indicate double bonds. PUBLISHERS RESOURCE GROUP

carbonate ($\text{ZnCO}_3 \cdot 2\text{Zn}(\text{OH})_2$). Leaching is a process by which soluble salts are dissolved out of soil, ore, or some other material. The basic zinc carbonate is decomposed by heat, yielding pure zinc oxide:



COMMON USES AND POTENTIAL HAZARDS

One of zinc oxide's most important uses historically is in paints and pigments. Common names by which it is now known, such as Chinese white and zinc white, are terms long used by artists for the compound. White paints and pigments made with zinc oxide retain their luster and purity longer than most other types of white paints, partly because zinc oxide does not react readily with hydrogen sulfide in air that causes most white paints to darken.

One of zinc oxide's most important properties is its ability to absorb ultraviolet (UV) light in sunlight. Because of this property, zinc oxide is often added to sunscreens and sunblocks to help protect a person from sunburn. The same property accounts for other important applications for zinc oxide, such as their use in rubber and plastic products. By absorbing ultraviolet light, zinc oxide protects the rubber or plastic from decomposing. The compound also has other applications in the rubber and plastic industries, including:

- In the curing of natural and synthetic rubber products, increasing the rate at which the product reaches its final chemical state;

Interesting Facts

- The term *philosopher's wool* dates to the period of alchemy when scholars often gave picturesque and descriptive names to the substances with which they worked. Philosopher's wool referred to the fluffy

white material that was formed when alchemists (also known as natural philosophers) heated zinc in air. The substance was also referred to as *nix alba*, or "white snow" for the same reason.

- As a fungicidal additive to rubber and plastic, preventing fungi from attacking and destroying products made from those materials;
- To increase the temperature at which rubber and plastic products remain stable and the amount of exposure to light they can withstand;
- To maintain the proper acidic properties of a product, reducing the rate at which they are likely to decay;
- To provide additional strength to the rubber or plastic product.

Besides its major uses in the rubber and plastic industries, zinc oxide has a number of applications in other fields, such as:

- As an additive in glass and ceramic materials, to provide greater heat resistance, greater resistance to breakage by shock, and high luster;
- In the manufacture of specialized types of sealants and adhesives;
- As a light-gathering agent in photocopy machines;
- As an additive in lubricants for the purpose of reducing wear and helping the lubricant withstand high pressures;
- In smokestacks as an aid in removing sulfur dioxide and other pollutant gases produced in factory operations; and

Words to Know

ALCHEMY An ancient field of study from which the modern science of chemistry evolved.

AQUEOUS SOLUTION A solution that consists of some material dissolved in water.

- In the manufacture of specialized packaging materials, especially those used for food products, because of the compound's ability to kill certain microorganisms that cause food decay.

In moderate amounts, zinc oxide is a relatively harmless compound. Exposure to zinc oxide dust may cause respiratory problems, such as coughing, upper respiratory tract irritation, chills, fever, nausea, and vomiting.

FOR FURTHER INFORMATION

“Application.” Nav Bharat Metallic Oxides Industries.

<http://www.navbharat.co.in/clients.htm> (accessed on November 19, 2005).

“Occupational Safety and Health Guideline for Zinc Oxide.” Occupational Safety & Health Administration.

<http://www.osha.gov/SLTC/healthguidelines/zincoxide/recognition.html> (accessed on November 19, 2005).

“Zinc Oxide.” Center for Advanced Microstructures and Devices, Louisiana State University.

http://www.camd.lsu.edu/msds/z/zinc_oxide.htm (accessed on November 19, 2005).

“Zinc Oxide Producers Association.”

<http://www.cefic.be/Templates/shwAssocDetails.asp?NID=5&HID=25&ID=172> (accessed on November 19, 2005).

appendices

Lists of compounds

Compounds by Formula	liii
Compounds by Element	lix
Compounds by Type	lxv

Compounds by Formula

AgI	Silver Iodide	CHCl_3	Chloroform
AgNO_3	Silver Nitrate	CH_2O	Formaldehyde
Ag_2O	Silver(I) Oxide	$-\text{[-CH}_2\text{C(CH}_3\text{)(COOH)CH}_2\text{-]}_n$	Polymethyl Methacrylate
Ag_2S	Silver(I) Sulfide	$\text{CH}_2=\text{C}(\text{CN})\text{COOCH}_3$	Cyanoacrylate
AlF_3	Aluminum Fluoride	$\text{CH}_2=\text{CHCH=CH}_2$	1,3-Butadiene
Al(OH)_3	Aluminum Hydroxide	$\text{CH}_2=\text{CHCH}_3$	Propylene
Al_2O_3	Aluminum Oxide	$\text{CH}_2=\text{CH(CH}_3\text{)CH=CH}_2$	Isoprene
CCl_2F_2	Dichlorodifluoromethane	$-\text{[-CH}_2\text{-CH(COONa)-]}_n$	Sodium Polyacrylate
CCl_4	Carbon Tetrachloride	$-\text{[CH}_2\text{CHC}_6\text{H}_5\text{-]}_n$	
$-\text{[-CF}_2\text{-]}_n$	Polytetrafluoroethylene	$-\text{[CH}_2\text{CH=CHCH}_2\text{-]}_n\text{-[CH}_2\text{CHC}_6\text{H}_5\text{-}}$	Poly(Styrene-Butadiene-Styrene)
$\text{CH}\equiv\text{CH}$	Acetylene	$-\text{[-CH}_2\text{CHCl-]}_n$	Polyvinyl Chloride
$-\text{[-CH(CH}_3\text{)CH}_2\text{-]}_n$	Polypropylene	$\text{CH}_2=\text{CH}_2$	Ethylene
		$(\text{CH}_2\text{CH}_2\text{Cl})_2\text{S}$	$2,2'$ -Dichlorodiethyl Sulfide

Compounds by Formula

$[-\text{CH}_2\text{CH}_2]_n$	CH_3OH	C_2H_4
Polyethylene	Methyl Alcohol	Ethylene
$[-\text{CH}_2\text{C}_6\text{H}_5]_n$	$(\text{CH}_3\text{O})(\text{OH})\text{C}_6\text{H}_3\text{CHO}$	$[\text{C}_2\text{H}_4]_n$
Polystyrene	Vanillin	Polyethylene
$\text{CH}_2\text{NO}_2\text{CHNOCH}_2\text{NO}_2$	CH_3SH	$\text{C}_2\text{H}_4\text{O}$
Nitroglycerin	Methyl Mercaptan	Ethylene Oxide
$\text{CH}_2\text{OHCHOHCH}_2\text{OH}$	$(\text{CH}_3)_2\text{CHCH}_2\text{CH}_2\text{NO}_2$	$\text{C}_2\text{H}_4\text{O}_2$
Glycerol	Amyl Nitrite	Acetic acid
$\text{CH}_2\text{OHCH}_2\text{OH}$	$\text{C}_5\text{H}_{11}\text{NO}_2$	$\text{C}_2\text{H}_6\text{O}$
Ethylene Glycol	Amyl Nitrite	Ethyl Alcohol
$(\text{CH}_2)_2\text{O}$	$(\text{CH}_3)_2\text{C}_5\text{H}_3\text{NSO}(\text{COOH})\text{NHCOR}$	$\text{C}_2\text{H}_6\text{O}_2$
Ethylene Oxide	Penicillin	Ethylene Glycol
$\text{CH}_3\text{CHOHCH}_3$	$(\text{CH}_3)_3\text{COCH}_3$	$[\text{C}_3\text{H}_3\text{O}_2\text{Na}]_n$
Isopropyl Alcohol	Methyl-t-butyl Ether	Sodium Polyacrylate
$\text{CH}_3\text{CHOHCOOH}$	CH_4	$\text{C}_3\text{H}_5\text{N}_3\text{O}_5$
Lactic Acid	Methane	Nitroglycerin
$\text{CH}_3\text{CH}_2\text{OH}$	CH_4S	$[\text{C}_3\text{H}_6]_n$
Ethyl Alcohol	Methyl Mercaptan	Polypropylene
CH_3COCH_3	CO	$\text{C}_3\text{H}_6\text{O}$
Dimethyl Ketone	Carbon Monoxide	Dimethyl Ketone
$\text{CH}_3\text{CONHC}_6\text{H}_4\text{OH}$	$[-\text{CO}(\text{CH}_2)_4\text{CO-NH}(\text{CH}_2)_6\text{NH-}]_n$	$\text{C}_3\text{H}_6\text{O}_3$
Acetaminophen	Nylon 6 and Nylon 66	Lactic Acid
$\text{CH}_3\text{COOCH}_2\text{CH}_2\text{CH}(\text{CH}_3)_2$	$[-\text{CO}(\text{CH}_2)_5\text{NH-}]_n$	C_3H_8
Isoamyl Acetate	Nylon 6 and Nylon 66	Propane
$\text{CH}_3\text{COOC}_2\text{H}_5$	$[-\text{CONH-C}_6\text{H}_4-\text{NCOO-CH}_2\text{CH}_2\text{-O-}]_n$	$\text{C}_3\text{H}_8\text{O}$
Ethyl Acetate	Polyurethane	Isopropyl Alcohol
$\text{CH}_3\text{COOC}_5\text{H}_{11}$	$\text{COOH}(\text{CH}_2)_2\text{CH}(\text{NH}_2)\text{COONa}$	$\text{C}_3\text{H}_8\text{O}_3$
Amyl Acetate	Monosodium Glutamate	Glycerol
$\text{CH}_3\text{COOC}_6\text{H}_4\text{COOH}$	CO_2	C_4H_6
Acetylsalicylic acid	Carbon Dioxide	1,3-Butadiene
CH_3COOH	C_2H_2	$\text{C}_4\text{H}_8\text{Cl}_2\text{S}$
Acetic acid	Acetylene	2,2'-Dichlorodiethyl Sulfide
$\text{CH}_3\text{C}_5\text{HN(OH)}(\text{CH}_2\text{OH})_2$	$\text{C}_2\text{H}_2\text{O}_4$	$\text{C}_4\text{H}_8\text{O}_2$
Pyridoxine	Oxalic Acid	Ethyl Acetate
$\text{CH}_3\text{C}_6\text{H}_9(\text{C}_3\text{H}_7)\text{OH}$	$[\text{C}_2\text{H}_3\text{Cl}]_n$	C_4H_{10}
Menthol	Polyvinyl Chloride	Butane

$C_4H_{10}S$	Butyl Mercaptan	C_6H_6	Benzene	$[C_7H_7]_n$	Polystyrene
$C_5H_4NC_4H_7NCH_3$	Nicotine	$C_6H_6Cl_6$	Gamma-1,2,3,4,5,6-Hexachlorocyclohexane	$C_8H_7N_3O_2$	Luminol
$C_5H_5NO_2$	Cyanoacrylate	C_6H_6O	Phenol	$C_7H_8N_4O_2$	Theobromine
C_5H_8	Isoprene	$[C_6H_7O_2(OH)_2OCS_2Na]_n$	Cellulose Xanthate	$[C_7H_{11}NaO_5S_2]_n$	Chloroform
$[C_5H_8]_n$	Polymethyl Methacrylate	$C_6H_8O_6$	Ascorbic Acid	$C_7H_{14}O_2$	Isoamyl Acetate Amyl Acetate
$C_5H_8NNaO_4$	Monosodium Glutamate	$C_6H_8O_7$	Citric Acid	C_8H_8	Styrene
$C_5H_{12}O$	Methyl-t-butyl Ether	$(C_6H_{10}O_5)_n$	Cellulose	$[C_8H_8]_n-[C_4H_6]_n-[C_8H_8]_n$	Poly(Styrene-Butadiene-Styrene)
$C_6H_2(CH_3)(NO_2)_3$	2,4,6-Trinitrotoluene	$[C_6H_{11}NO]_n$	Nylon 6 and Nylon 66	$C_8H_9NO_2$	Acetaminophen
$C_6H_3ClOH-O-C_6H_3Cl_2$	Triclosan	$C_6H_{11}NHSO_3Na$	Sodium Cyclamate	C_8H_{10}	Ethylbenzene
$C_6H_3Cl_2NHCONHC_6H_4Cl$	Triclocarban	$C_6H_{12}NNaSO_3$	Sodium Cyclamate	$C_8H_{10}N_4O_2$	Caffeine
$C_6H_5CH=CHCHO$	Cinnamaldehyde	$C_6H_{12}O_2$	Butyl Acetate	$C_8H_{11}NO_3$	Pyridoxine
$C_6H_5CH=CH_2$	Styrene	$C_6H_{12}O_6$	Fructose Glucose	C_9H_8O	Cinnamaldehyde
$C_6H_5CH(CH_3)_2$	Cumene	C_6H_{14}	Hexane	$C_9H_8O_4$	Acetylsalicylic acid
$C_6H_5CH_3$	Toluene	$C_6H_4(CH_3)CON(C_2H_5)_2$	N,N-Diethyl-3-Methylbenzamide	C_9H_{12}	Cumene
C_6H_5COOH	Benzoinic Acid	$C_7H_5NO_3S$	Saccharin	$C_{10}H_8$	Naphthalene
$C_6H_5C_2H_5$	Ethylbenzene	$C_7H_5N_3O_6$	2,4,6-Trinitrotoluene	$C_{10}H_{14}N_2$	Nicotine
$C_6H_5NO_2$	Niacin	$C_7H_6O_2$	Benzoic Acid	$C_{10}H_{16}O$	Camphor
C_6H_5OH	Phenol				

Compounds by Formula

$C_{10}H_{20}O$	$C_{16}H_{19}N_3O_5S$	CaO
Menthol	Amoxicillin	Calcium Oxide
$C_{11}H_{16}O_2$	$C_{17}H_{20}N_4O_6$	$Ca(OH)_2$
Butylated Hydroxyanisole and Butylated Hydroxyto- luene (BHA)	Riboflavin	Calcium Hydroxide
$C_{12}H_7Cl_3O_2$	$C_{19}H_{19}N_7O_6$	$CaSO_4$
Triclosan	Folic Acid	Calcium Sulfate
$C_{12}H_{16}N_4O_{18}$	$C_{19}H_{28}O_2$	$CaSiO_3$
Cellulose Nitrate	Testosterone	Calcium Silicate
$C_{12}H_{17}ClN_4OS$	$C_{20}H_{30}O$	$Ca(H_2PO_4)_2$
Thiamine	Retinol	Calcium Phosphate
$C_{12}H_{17}NO$	$C_{27}H_{45}OH$	$Ca_3(PO_4)_2$
N,N-Diethyl-3-Methylbenza- mide	Cholesterol	Calcium Phosphate
$[C_{12}H_{22}N_2O_2]_n$	$C_{28}H_{34}N_2O_3$	$(ClC_6H_4)_2CHCCl_3$
Nylon 6 and Nylon 66	Denatonium Benzoate	Dichlorodiphenyltrichloro- ethane
$C_{12}H_{22}O_{11}$	$C_{29}H_{50}O$	$-ClO_4$
Lactose	Alpha-Tocopherol	Perchlorates
Sucrose	$C_{35}H_{28}O_5N_4Mg$	CuO
$C_{13}H_9Cl_3N_2O$	Chlorophyll	Copper(II) Oxide
Triclocarban	$C_{35}H_{30}O_5N_4Mg$	$CuSO_4$
$C_{13}H_{18}O_2$	Beta-Carotene	Copper(II) Sulfate
2-(4-Isobutylphenyl)propionic Acid	$C_{54}H_{70}O_6N_4Mg$	Cu_2O
$C_{14}H_9Cl_5$	Chlorophyll	Copper(I) Oxide
Dichlorodiphenyltrichloro- ethane	$C_{55}H_{70}O_6N_4Mg$	FeO
$C_{14}H_{14}O_3$	Chlorophyll	Iron(II) Oxide
Naproxen	$C_{55}H_{72}O_5N_4Mg$	Fe_2O_3
$C_{14}H_{18}N_2O_5$	Cyanocobalamin	Iron(III) Oxide
L-Aspartyl-L-Phenylalanine Methyl Ester	$C_{76}H_{52}O_{46}$	$HCHO$
$C_{15}H_{24}O$	Tannic Acid	Formaldehyde
Butylated Hydroxyanisole and Butylated Hydroxyto- luene (BHT)	$CaCO_3$	HCl
$C_{16}H_{18}N_2OS$	Calcium Carbonate	Hydrogen Chloride
Penicillin	$CaHPO_4$	HNO_3
	Calcium Phosphate	Nitric Acid
		HOH
		Water

HOOCH₂C(OH)(COOH)CH₂COOH	K₂CO₃ K₃₄NO₁₅	NaCl
Citric Acid	Potassium Carbonate	Sodium Chloride
HOOCCOOH	K₂SO₄	NaClO
Oxalic Acid	Potassium Sulfate	Sodium Hypochlorite
H₂O	MgCl₂	NaF
Water	Magnesium Chloride	Sodium Fluoride
H₂O₂	MgO	NaHCO₃
Hydrogen Peroxide	Magnesium Oxide	Sodium Bicarbonate
H₂SO₄	Mg(OH)₂	NaH₂PO₄
Sulfuric Acid	Magnesium Hydroxide	Sodium Phosphate
H₃BO₃	MgSO₄	NaOH
Boric Acid	Magnesium Sulfate	Sodium Hydroxide
H₃PO₄	Mg₃Si₄O₁₀(OH)₂	Na₂B₄O₇
Phosphoric Acid	Magnesium Silicate Hydroxide	Sodium Tetraborate
HgS	(NH₂)₂CO	Na₂B₄O₇·10H₂O
Mercury(II) Sulfide	Urea	Sodium Tetraborate
KAl(SO₄)₂	NH₃	Na₂CO₃
Aluminum Potassium Sulfate	Ammonia	Sodium Carbonate
KCl	(NH₄)₂SO₄	Na₂HPO₄
Potassium Chloride	Ammonium Sulfate	Sodium Phosphate
KF	NH₄Cl	Na₂SO₃
Potassium Fluoride	Ammonium Chloride	Sodium Sulfite
KHCO₃	NH₄NO₃	Na₂S₂O₃
Potassium Bicarbonate	Ammonium Nitrate	Sodium Thiosulfate
KHC₄H₄O₆	NH₄OH	Na₂SiO₃
Potassium Bitartrate	Ammonium Hydroxide	Sodium Silicate
KHSO₄	NO	Na₃PO₄
Potassium Bisulfate	Nitric Oxide	Sodium Phosphate
KI	NO₂	SO₂
Potassium Iodide	Nitrogen Dioxide	Sulfur Dioxide
KNO₃	N₂O	SiO₂
Potassium Nitrate	Nitrous Oxide	Silicon Dioxide
KOH	NaBO₃	SnF₂
Potassium Hydroxide	Sodium Perborate	Stannous Fluoride
	NaC₂H₃O₂	ZnO
	Sodium Acetate	Zinc Oxide

Compounds by Element

ALUMINUM

Aluminum Fluoride
Aluminum Hydroxide
Aluminum Oxide
Aluminum Potassium Sulfate

Amoxicillin
Amyl Acetate
Amyl Nitrite
Ascorbic Acid
Benzene
Benzoic Acid
Beta-Carotene
Butane

BORON

Boric Acid
Sodium Perborate
Sodium Tetraborate

Butyl Acetate
Butyl Mercaptan
Butylated Hydroxyanisole
and Butylated
Hydroxytoluene

CALCIUM

Calcium Carbonate
Calcium Hydroxide
Calcium Oxide
Calcium Phosphate
Calcium Silicate
Calcium Sulfate

Caffeine
Calcium Carbonate
Camphor
Carbon Dioxide
Carbon Monoxide
Carbon Tetrachloride
Cellulose

CARBON

1,3-Butadiene
2-(4-Isobutylphenyl)propionic
Acid
2,2'-Dichlorodiethyl
Sulfide
2,4,6-Trinitrotoluene
Acetaminophen
Acetic acid
Acetylene
Acetylsalicylic acid
Alpha-Tocopherol

Cellulose Nitrate
Cellulose Xanthate
Chloroform
Chlorophyll
Cholesterol
Cinnamaldehyde
Citric Acid
Collagen
Cumene
Cyanoacrylate
Cyanocobalamin
Denatonium Benzoate
Dichlorodifluoromethane

Compounds by Element

Dichlorodiphenyltrichloroethane	Polycarbonates	Magnesium Chloride
Dimethyl Ketone	Polyethylene	Perchlorates
Ethyl Acetate	Polymethyl Methacrylate	Polyvinyl Chloride
Ethyl Alcohol	Polypropylene	Potassium Chloride
Ethylbenzene	Polysiloxane	Sodium Chloride
Ethylene	Polystyrene	Sodium Hypochlorite
Ethylene Glycol	Polytetrafluoroethylene	Thiamine
Ethylene Oxide	Polyurethane	Triclocarban
Folic Acid	Polyvinyl Chloride	Triclosan
Formaldehyde	Potassium Bicarbonate	
Fructose	Potassium Bitartrate	
Gamma-1,2,3,4,5,6-Hexachlorocyclohexane	Potassium Carbonate	
Gelatin	Propane	COBALT
Glucose	Propylene	Cyanocobalamin
Glycerol	Pyridoxine	
Hexane	Retinol	COPPER
Isoamyl Acetate	Riboflavin	Copper(I) Oxide
Isoprene	Saccharin	Copper(II) Oxide
Isopropyl Alcohol	Sodium Acetate	Copper(II) Sulfate
Lactic Acid	Sodium Bicarbonate	
Lactose	Sodium Carbonate	FLUORINE
L-Aspartyl-L-Phenylalanine Methyl Ester	Sodium Cyclamate	Aluminum Fluoride
Luminol	Sodium Polyacrylate	Dichlorodifluoromethane
Menthol	Styrene	Polytetrafluoroethylene
Methane	Sucrose	Potassium Fluoride
Methyl Alcohol	Sucrose Polyester	Sodium Fluoride
Methyl Mercaptan	Tannic Acid	Stannous Fluoride
Methyl-t-butyl Ether	Testosterone	
Monosodium Glutamate	Theobromine	HYDROGEN
N,N-Diethyl-3-Methylbenzamide	Thiamine	1,3-Butadiene
Naphthalene	Toluene	2-(4-Isobutylphenyl)propionic Acid
Naproxen	Triclocarban	2,2'-Dichlorodiethyl Sulfide
Niacin	Triclosan	2,4,6-Trinitrotoluene
Nicotine	Urea	Acetaminophen
Nitroglycerin	Vanillin	Acetic acid
Nylon 6 and Nylon 66		Acetylene
Oxalic Acid	CHLORINE	Acetylsalicylic acid
Pectin	2,2'-Dichlorodiethyl Sulfide	Alpha-Tocopherol
Penicillin	Ammonium Chloride	Ammonia
Petrolatum	Carbon Tetrachloride	Ammonium Chloride
Petroleum	Chloroform	Ammonium Hydroxide
Phenol	Dichlorodifluoromethane	Ammonium Nitrate
Poly(Styrene-Butadiene-Styrene)	Dichlorodiphenyltrichloroethane	Ammonium Sulfate
	Gamma-1,2,3,4,5,6-Hexachlorocyclohexane	Amoxicillin
	Hydrogen Chloride	Amyl Acetate
		Amyl Nitrite
		Ascorbic Acid

Benzene	Lactic Acid	Retinol
Benzoic Acid	Lactose	Riboflavin
Beta-Carotene	L-Aspartyl-L-Phenylalanine	Saccharin
Boric Acid	Methyl Ester	Sodium Acetate
Butane	Luminol	Sodium Bicarbonate
Butyl Acetate	Magnesium Hydroxide	Sodium Cyclamate
Butyl Mercaptan	Magnesium Silicate	Sodium Hydroxide
Butylated Hydroxyanisole and Butylated Hydroxytoluene	Hydroxide	Sodium Polyacrylate
Caffeine	Menthol	Styrene
Calcium Hydroxide	Methane	Sucrose
Calcium Phosphate	Methyl Alcohol	Sucrose Polyester
Camphor	Methyl Mercaptan	Sulfuric Acid
Cellulose	Methyl-t-butyl Ether	Tannic Acid
Cellulose Nitrate	Monosodium Glutamate	Testosterone
Cellulose Xanthate	N,N-Diethyl-3-Methyl- benzamide	Theobromine
Chloroform	Naphthalene	Thiamine
Chlorophyll	Naproxen	Toluene
Cholesterol	Niacin	Triclocarban
Cinnamaldehyde	Nicotine	Triclosan
Citric Acid	Nitric Acid	Urea
Collagen	Nitroglycerin	Vanillin
Cumene	Nylon 6 and Nylon 66	Water
Cyanoacrylate	Oxalic Acid	
Cyanocobalamin	Pectin	IODINE
Denatonium Benzoate	Penicillin	Potassium Iodide
Dichlorodiphenyltrichloro- ethane	Petrolatum	Silver Iodide
Dimethyl Ketone	Petroleum	
Ethyl Acetate	Phenol	IRON
Ethyl Alcohol	Phosphoric Acid	Iron(II) Oxide
Ethylbenzene	Poly(Styrene-Butadiene- Styrene)	Iron(III) Oxide
Ethylene	Polycarbonates	
Ethylene Glycol	Polyethylene	MAGNESIUM
Ethylene Oxide	Polymethyl	Chlorophyll
Folic Acid	Methacrylate	Magnesium Chloride
Formaldehyde	Polypropylene	Magnesium Hydroxide
Fructose	Polysiloxane	Magnesium Oxide
Gamma-1,2,3,4,5,6-Hexachloro- cyclohexane	Polystyrene	Magnesium Silicate
Gelatin	Polyurethane	Hydroxide
Glucose	Polyvinyl Chloride	Magnesium Sulfate
Glycerol	Potassium Bicarbonate	
Hexane	Potassium Bisulfate	MERCURY
Hydrogen Chloride	Potassium Bitartrate	Mercury(II) Sulfide
Isoamyl Acetate	Potassium Hydroxide	
Isoprene	Propane	NITROGEN
Isopropyl Alcohol	Propylene	2,4,6-Trinitrotoluene
	Pyridoxine	Acetaminophen

Compounds by Element

Ammonia	Acetic acid	Folic Acid
Ammonium Chloride	Acetylsalicylic acid	Formaldehyde
Ammonium Hydroxide	Alpha-Tocopherol	Fructose
Ammonium Nitrate	Aluminum Hydroxide	Gelatin
Ammonium Sulfate	Aluminum Potassium Sulfate	Glucose
Amoxicillin	Aluminum Oxide	Glycerol
Amyl Nitrite	Ammonium Hydroxide	Hydrogen Peroxide
Caffeine	Ammonium Nitrate	Iron(II) Oxide
Cellulose Nitrate	Ammonium Sulfate	Iron(III) Oxide
Chlorophyll	Amoxicillin	Isoamyl Acetate
Collagen	Amyl Acetate	Isopropyl Alcohol
Cyanoacrylate	Amyl Nitrite	Lactic Acid
Cyanocobalamin	Ascorbic Acid	Lactose
Denatonium Benzoate	Benzoic Acid	L-Aspartyl-L-Phenylalanine
Folic Acid	Boric Acid	Methyl Ester
Gelatin	Butyl Acetate	Luminol
L-Aspartyl-L-Phenylalanine	Butylated Hydroxyanisole	Magnesium Hydroxide
Methyl Ester	and Butylated Hydro-	Magnesium Oxide
Luminol	xytoluene	Magnesium Silicate
Monosodium Glutamate	Caffeine	Hydroxide
N,N-Diethyl-3-Methyl-	Calcium Carbonate	Magnesium Sulfate
benzamide	Calcium Hydroxide	Menthol
Niacin	Calcium Oxide	Methyl Alcohol
Nicotine	Calcium Phosphate	Methyl-t-butyl Ether
Nitric Acid	Calcium Silicate	Monosodium Glutamate
Nitric Oxide	Calcium Sulfate	N,N-Diethyl-3-Methyl-
Nitrogen Dioxide	Camphor	benzamide
Nitroglycerin	Carbon Dioxide	Naproxen
Nylon 6 and Nylon 66	Carbon Monoxide	Niacin
Penicillin	Cellulose	Nitric Acid
Polyurethane	Cellulose Nitrate	Nitric Oxide
Potassium Nitrate	Cellulose Xanthate	Nitrogen Dioxide
Pyridoxine	Chlorophyll	Nitroglycerin
Riboflavin	Cholesterol	Nitrous Oxide
Saccharin	Cinnamaldehyde	Nylon 6 and Nylon 66
Silver Nitrate	Citric Acid	Oxalic Acid
Sodium Cyclamate	Collagen	Pectin
Theobromine	Copper(I) Oxide	Penicillin
Thiamine	Copper(II) Oxide	Perchlorates
Triclocarban	Copper(II) Sulfate	Petroleum
Urea	Cyanoacrylate	Phenol
Nitrous Oxide	Cyanocobalamin	Phosphoric Acid
OXYGEN	Denatonium Benzoate	Polycarbonates
2-(4-Isobutylphenyl)propionic	Dimethyl Ketone	Polymethyl Methacrylate
Acid	Ethyl Acetate	Polysiloxane
2,4,6-Trinitrotoluene	Ethyl Alcohol	Polyurethane
Acetaminophen	Ethylene Glycol	Potassium Bicarbonate
	Ethylene Oxide	Potassium Bisulfate

Potassium Bitartrate
 Potassium Carbonate
 Potassium Hydroxide
 Potassium Nitrate
 Potassium Sulfate
 Pyridoxine
 Retinol
 Riboflavin
 Saccharin
 Silicon Dioxide
 Silver Nitrate
 Silver(I) Oxide
 Sodium Acetate
 Sodium Bicarbonate
 Sodium Carbonate
 Sodium Cyclamate
 Sodium Hydroxide
 Sodium Hypochlorite
 Sodium Perborate
 Sodium Phosphate
 Sodium Polyacrylate
 Sodium Silicate
 Sodium Sulfite
 Sodium Tetraborate
 Sodium Thiosulfate
 Sucrose
 Sucrose Polyester
 Sulfur Dioxide
 Sulfuric Acid
 Tannic Acid
 Testosterone
 Theobromine
 Thiamine
 Triclocarban
 Triclosan
 Urea
 Vanillin
 Water
 Zinc Oxide

PHOSPHORUS

Calcium Phosphate

Phosphoric Acid
 Sodium Phosphate

POTASSIUM

Aluminum Potassium Sulfate
 Potassium Bicarbonate
 Potassium Bisulfate
 Potassium Bitartrate
 Potassium Carbonate
 Potassium Chloride
 Potassium Fluoride
 Potassium Hydroxide
 Potassium Iodide
 Potassium Nitrate
 Potassium Sulfate

SILICON

Calcium Silicate
 Magnesium Silicate Hydroxide
 Polysiloxane
 Silicon Dioxide
 Sodium Silicate

SILVER

Silver Iodide
 Silver Nitrate
 Silver(I) Oxide
 Silver(I) Sulfide

SODIUM

Cellulose Xanthate
 Monosodium Glutamate
 Sodium Acetate
 Sodium Bicarbonate
 Sodium Carbonate
 Sodium Chloride
 Sodium Fluoride
 Sodium Hydroxide

Sodium Hypochlorite
 Sodium Perborate
 Sodium Phosphate
 Sodium Polyacrylate
 Sodium Silicate
 Sodium Sulfite
 Sodium Tetraborate
 Sodium Thiosulfate

SULFUR

2,2'-Dichlorodiethyl Sulfide
 Aluminum Potassium Sulfate
 Ammonium Sulfate
 Amoxicillin
 Butyl Mercaptan
 Calcium Sulfate
 Cellulose Xanthate
 Copper(II) Sulfate
 Magnesium Sulfate
 Mercury(II) Sulfide
 Methyl Mercaptan
 Penicillin
 Potassium Bisulfate
 Potassium Sulfate
 Saccharin
 Silver(I) Sulfide
 Sodium Cyclamate
 Sodium Sulfite
 Sodium Thiosulfate
 Sulfur Dioxide
 Sulfuric Acid
 Thiamine

TIN
 Stannous Fluoride

ZINC
 Zinc Oxide

Compounds by Type

ACID

2-(4-Isobutylphenyl)propionic Acid
Acetic acid
Acetylsalicylic acid
Ascorbic Acid
Benzoic Acid
Boric Acid
Butyl Acetate
Citric Acid
Denatonium Benzoate
Folic Acid
Hydrogen Chloride
Lactic Acid
Naproxen
Niacin
Nitric Acid
Oxalic Acid
Penicillin
Phosphoric Acid
Potassium Bicarbonate
Potassium Bisulfate
Potassium Bitartrate
Sodium Bicarbonate
Sulfuric Acid
Tannic Acid

ALCOHOL

Ethyl Alcohol
Ethylene Glycol
Glycerol
Isopropyl Alcohol

Menthol

Methyl Alcohol
Retinol

ALDEHYDE

Cinnamaldehyde
Formaldehyde

ALKALOID

Caffeine
Nicotine
Theobromine

ALKANE

Butane
Hexane
Methane
Propane

ALKENE

1,3-Butadiene
Ethylene
Propylene

ALKYNE

Acetylene

AMIDE

Acetaminophen

Compounds by Type

BASE

Aluminum Hydroxide
Ammonia
Ammonium Hydroxide
Caffeine
Calcium Hydroxide
Magnesium Hydroxide
Potassium Hydroxide
Sodium Hydroxide
Theobromine

CARBOHYDRATE

Cellulose
Cellulose Nitrate
Fructose
Glucose
Lactose
Sucrose

CARBOXYLIC ACID

Acetic acid
Acetylsalicylic acid
Butyl Acetate
Citric Acid
Lactic Acid
Naproxen
Niacin
Oxalic Acid

ESTER

Amyl Acetate
Amyl Nitrite
Cyanoacrylate
Ethyl Acetate
Isoamyl Acetate
L-Aspartyl-L-Phenylalanine
Methyl Ester
Nitroglycerin

ETHER

Ethylene Oxide
Methyl-t-butyl Ether
Vanillin

HYDROCARBON

1,3-Butadiene
Acetylene
Benzene
Beta-Carotene
Butane
Cumene
Ethylbenzene
Ethylene
Hexane
Isoprene
Methane
Naphthalene
Propane
Propylene
Styrene
Toluene

INORGANIC

Aluminum Fluoride
Aluminum Hydroxide
Aluminum Oxide
Aluminum Potassium Sulfate
Ammonia
Ammonium Chloride
Ammonium Hydroxide
Ammonium Nitrate
Ammonium Sulfate
Boric Acid
Calcium Carbonate
Calcium Hydroxide
Calcium Oxide
Calcium Phosphate
Calcium Silicate
Calcium Sulfate
Carbon Dioxide
Carbon Monoxide
Copper(I) Oxide
Copper(II) Oxide
Copper(II) Sulfate
Hydrogen Chloride
Iron(II) Oxide
Iron(III) Oxide
Magnesium Chloride
Magnesium Hydroxide
Magnesium Oxide

Magnesium Silicate Hydroxide

Magnesium Sulfate

Mercury(II) Sulfide

Nitric Acid

Nitric Oxide

Nitrogen Dioxide

Nitrous Oxide

Perchlorates

Phosphoric Acid

Polysiloxane

Potassium Bicarbonate

Potassium Bisulfate

Potassium Bitartrate

Potassium Carbonate

Potassium Chloride

Potassium Fluoride

Potassium Hydroxide

Potassium Iodide

Potassium Nitrate

Potassium Sulfate

Silicon Dioxide

Silver Iodide

Silver Nitrate

Silver(I) Oxide

Silver(I) Sulfide

Sodium Acetate

Sodium Bicarbonate

Sodium Carbonate

Sodium Chloride

Sodium Fluoride

Sodium Hydroxide

Sodium Hypochlorite

Sodium Perborate

Sodium Phosphate

Sodium Silicate

Sodium Sulfite

Sodium Tetraborate

Sodium Thiosulfate

Stannous Fluoride

Sulfur Dioxide

Sulfuric Acid

Water

Zinc Oxide

KETONE

Camphor

Dimethyl Ketone

METALLIC OXIDE

Aluminum Oxide
Calcium Oxide
Copper(I) Oxide
Copper(II) Oxide
Iron(II) Oxide
Iron(III) Oxide
Magnesium Oxide
Silver(I) Oxide
Zinc Oxide

NONMETALLIC OXIDE

Carbon Dioxide
Carbon Monoxide
Hydrogen Peroxide
Nitric Oxide
Nitrogen Dioxide
Nitrous Oxide
Silicon Dioxide
Sulfur Dioxide

ORGANIC

1,3-Butadiene
2-(4-Isobutylphenyl)propionic Acid
2,2'-Dichlorodiethyl Sulfide
2,4,6-Trinitrotoluene
Acetaminophen
Acetic acid
Acetylene
Acetylsalicylic acid
Alpha-Tocopherol
Amoxicillin
Amyl Acetate
Amyl Nitrite
Ascorbic Acid
Benzene
Benzoic Acid
Beta-Carotene
Butane
Butyl Acetate
Butyl Mercaptan
Butylated Hydroxyanisole and Butylated Hydroxytoluene
Caffeine
Camphor

Carbon Tetrachloride

Cellulose
Cellulose Nitrate
Cellulose Xanthate

Chloroform
Chlorophyll
Cholesterol
Cinnamaldehyde
Citric Acid
Collagen

Cumene
Cyanoacrylate
Cyanocobalamin

Denatonium Benzoate
Dichlorodifluoromethane
Dichlorodiphenyltrichloroethane
Dimethyl Ketone

Ethyl Acetate
Ethyl Alcohol
Ethylbenzene
Ethylene

Ethylene Glycol
Ethylene Oxide

Folic Acid
Formaldehyde
Fructose
Gamma-1,2,3,4,5,6-Hexachlorocyclohexane

Glucose
Glycerol
Hexane

Hydrogen Peroxide
Isoamyl Acetate
Isoprene
Isopropyl Alcohol

Lactic Acid
Lactose
L-Aspartyl-L-Phenylalanine Methyl Ester

Luminol
Menthol
Methane
Methyl Alcohol
Methyl Mercaptan
Methyl-t-butyl Ether
Monosodium Glutamate

N,N-Diethyl-3-Methylbenzamide

Naphthalene

Naproxen

Niacin

Nicotine

Nitroglycerin

Nylon 6 and Nylon 66

Oxalic Acid

Penicillin

Phenol

Poly(Styrene-Butadiene-Styrene)

Polycarbonates

Polyethylene

Polymethyl Methacrylate

Polypropylene

Polystyrene

Polytetrafluoroethylene

Polyurethane

Polyvinyl Chloride

Propane

Propylene

Pyridoxine

Retinol

Riboflavin

Saccharin

Sodium Cyclamate

Sodium Polyacrylate

Styrene

Sucrose

Sucrose Polyester

Tannic Acid

Testosterone

Theobromine

Thiamine

Toluene

Triclocarban

Triclosan

Urea

Vanillin

PHENOL

Butylated Hydroxyanisole and Butylated Hydroxytoluene

Phenol

Compounds by Type

POLYMER

Cellulose
Cellulose Nitrate
Cellulose Xanthate
Nylon 6 and Nylon 66
Poly(Styrene-Butadiene-Styrene)
Polycarbonates
Polyethylene
Polymethyl Methacrylate
Polypropylene
Polysiloxane
Polystyrene
Polytetrafluoroethylene
Polyurethane
Polyvinyl Chloride
Sodium Polyacrylate

SALT

Aluminum Fluoride
Aluminum Potassium Sulfate
Ammonium Chloride
Ammonium Nitrate
Ammonium Sulfate

Calcium Carbonate
Calcium Phosphate
Calcium Silicate
Calcium Sulfate
Copper(II) Sulfate
Magnesium Chloride
Magnesium Silicate Hydroxide
Magnesium Sulfate
Mercury(II) Sulfide
Monosodium Glutamate
Perchlorates
Potassium Bicarbonate
Potassium Bisulfate
Potassium Bitartrate
Potassium Carbonate
Potassium Chloride
Potassium Fluoride
Potassium Iodide
Potassium Nitrate
Potassium Sulfate
Silver Iodide
Silver Nitrate
Silver(I) Sulfide
Sodium Acetate

Sodium Bicarbonate
Sodium Carbonate
Sodium Chloride
Sodium Cyclamate
Sodium Fluoride
Sodium Hypochlorite
Sodium Perborate
Sodium Phosphate
Sodium Silicate
Sodium Sulfite
Sodium Tetraborate
Sodium Thiosulfate
Stannous Fluoride

VITAMIN

Alpha-Tocopherol
Ascorbic Acid
Cyanocobalamin
Folic Acid
Niacin
Pyridoxine
Retinol
Riboflavin
Thiamine

for further information

BOOKS

- Attenborough, David. *The Private Life of Plants*. Princeton, NJ: Princeton University Press, 1995.
- Brody, Tom. *Nutritional Biochemistry*. San Diego: Academic Press, 1998.
- Buchanan, B. B., W. Gruisse, and R. L. Jones. *Biochemistry and Molecular Biology of Plants*. Rockville, MD: American Society of Plant Physiologists, 2000.
- Buechel, K. H., et al. *Industrial Inorganic Chemistry*. New York: VCH, 2000.
- Buschmann, Helmut, et al. *Analgesics: From Chemistry and Pharmacology to Clinical Application*. New York: Wiley VCH, 2002.
- Butler, A. R., and R. Nicholson. *Life, Death and Nitric Oxide*. London: Royal Society of Chemistry, 2003.
- Cagin, Seth, and Philip Dray. *Between Earth and Sky: How CFCs Changed Our World and Endangered the Ozone Layer*. New York: Pantheon Books, 1993.
- Carpenter, Kenneth J. *The History of Scurvy and Vitamin C*. Cambridge, UK: Cambridge University Press, 1986.
- Carson, Rachel. *Silent Spring*. Boston: Houghton Mifflin, 1962.

For Further Information

- Cavitch, Susan Miller. *The Natural Soap Book: Making Herbal and Vegetable-Based Soaps*. Markham, Canada: Storey Publishing, 1995.
- Challem, Jack, and Melissa Diane Smith. *Basic Health Publications User's Guide to Vitamin E: Don't Be a Dummy: Become an Expert on What Vitamin E Can Do for Your Health*. North Bergen, NJ: Basic Health Publications, 2002.
- Chalmers, Louis. *Household and Industrial Chemical Specialties*. Vol. 1. New York: Chemical Publishing Co., Inc., 1978.
- Cherniske, Stephen. *Caffeine Blues: Wake Up to the Hidden Dangers of America's #1 Drug*. New York: Warner Books, 1998.
- "Cholesterol, Other Lipids, and Lipoproteins." In *In-Depth Report*. Edited by Julia Goldrosen. Atlanta: A.D.A.M., 2004.
- Cooper, P. W., and S. R. Kurowski. *Introduction to the Technology of Explosives*. New York: Wiley VCH, 1997.
- Cornell, Rochelle M., and Udo Schwertmann. *The Iron Oxides: Structure, Properties, Reactions, and Uses*, 2nd ed. New York: Wiley VCH, 2003.
- CRC Handbook of Chemistry and Physics*. David R. Lide, editor in chief. 86th ed. Boca Raton, FL: Taylor & Francis, 2005.
- Dean, Carolyn. *The Miracle of Magnesium*. New York: Ballantine Books, 2003.
- Dick, John S., and R. A. Annicelli, eds. *Rubber Technology: Compounding and Testing for Performance*. Cincinnati, OH: Hanser Gardner Publications, 2001.
- Dunlap, Thomas. *DDT: Scientists, Citizens, and Public Policy*. Princeton, NJ: Princeton University Press, 1983.
- Dwyer, Bob, et al. *Carbon Monoxide: A Clear and Present Danger*. Mount Prospect, IL: ESCO Press, 2004.
- Eades, Mary Dan. *The Doctor's Complete Guide to Vitamins and Minerals*. New York: Dell, 2000.
- Environment Canada, Health Canada. *Ethylene Oxide*. Ottawa: Environment Canada, 2001.
- Food Antioxidants: Technological, Toxicological, and Health Perspectives*, D. L. Madhavi, S. S. Deshpande, and D. K. Salunkhe, eds. New York: Dekker, 1996.
- Gahlinger, Paul M. *Illegal Drugs: A Complete Guide to Their History*. Salt Lake City, UT: Sagebrush Press, 2001.
- Genge, Ngaire E. *The Forensic Casebook: The Science of Crime Scene Investigation*. New York: Ballantine, 2002.
- Grimm, Tom, and Michele Grimm. *The Basic Book of Photography*. New York: Plume Books, 2003.

- Harte, John, et al. *Toxics A to Z*. Berkeley: University of California Press, 1991.
- Hermes, Matthew E. *Enough for One Lifetime: Wallace Carothers, Inventor of Nylon*. Philadelphia: Chemical Heritage Foundation, 1996.
- Holgate, S. T., et al. *Air Pollution and Health*. New York: Academic Press, 1999.
- Hyne, Norman J. *Nontechnical Guide to Petroleum Geology, Exploration, Drilling and Production*, 2nd ed. Tulsa, OK: Pennwell Books, 2001.
- Jeffreys, Diarmuid. *Aspirin: The Remarkable Story of a Wonder Drug*. New York: Bloomsbury, 2004.
- Johnson, Peter S. *Rubber Processing: An Introduction*. Cincinnati, OH: Hanser Gardner Publications, 2001.
- Karger Kocsis, J. *Polypropylene—An A-Z Reference*. New York: Springer Verlag, 1998.
- Kirk-Othmer Encyclopedia of Chemical Technology*, 4th ed. New York: John Wiley & Sons, 1991.
- Knox, J. Paul, and Graham B. Seymour, eds. *Pectins and Their Manipulation*. Boca Raton, FL: CFC Press, June 2002.
- Kurlansky, Mark. *Salt: A World History*. New York: Walker, 2002.
- Leffler, William L. *Petroleum Refining in Nontechnical Language*. Tulsa, OK: Pennwell Books, 2000.
- Mead, Clifford, and Thomas Hager, eds. *Linus Pauling: Scientist and Peacemaker*. Portland, OR: Oregon State University Press, 2001.
- Mebane, Robert C., and Thomas R. Rybolt. *Plastics and Polymers*. New York: Twenty First Century, 1995.
- Meikle, Jeffrey L. *American Plastic: A Cultural History*. New Brunswick, NJ: Rutgers University Press, 1995.
- Menhard, Francha Roffe. *The Facts about Inhalants*. New York: Benchmark Books, 2004.
- Misra, Chanakya. *Industrial Alumina Chemicals*. Washington, DC: American Chemical Society, 1986.
- Mosby's Medical, Nursing, and Allied Health Dictionary*, 5th ed. St. Louis: Mosby, 1998.
- Multhauf, Robert P., and Christine M. Roane. "Nitrates." In *Dictionary of American History*. 3rd ed., vol. 6. Stanley I. Kutler, ed. New York: Charles Scribner's Sons, 2003.
- Nabors, Lyn O'Brien, ed. *Alternative Sweeteners (Food Science and Technology)*, 3rd rev. London: Marcel Dekker, 2001.

For Further Information

- Packer, Lester, and Carol Colman. *The Antioxidant Miracle: Put Lipoic Acid, Pycogenol, and Vitamins E and C to Work for You.* New York: Wiley, 1999.
- Patnaik, Pradyot. *Handbook of Inorganic Chemicals.* New York: McGraw Hill, 2003.
- Rain, Patricia. *Vanilla: A Cultural History of the World's Most Popular Flavor and Fragrance.* Edited by Jeremy P. Tarcher. New York: Penguin Group USA, 2004.
- Rainsford, K. D., ed. *Ibuprofen: A Critical Bibliographic Review.* Bethesda, MD: CCR Press, 1999.
- Richardson, H. W., ed. *Handbook of Copper Compounds and Applications.* New York: Marcel Dekker, 1997.
- Sherman, Josepha, and Steve Brick. *Fossil Fuel Power.* Mankato, MN: Capstone Press, 2003.
- Snyder, C. H. *The Extraordinary Chemistry of Ordinary Things,* 4th ed. New York: John Wiley and Sons, 2002.
- Strange, Veronica. *The Meaning of Water.* Oxford, UK: Berg Publishers, 2004.
- Stratmann, Linda. *Chloroform: The Quest for Oblivion.* Phoenix Mill, UK: Sutton Publishing Co., 2003.
- Tegethoff, F. Wolfganga, with Johannes Rohleder and Evelyn Kroker, eds. *Calcium Carbonate: From the Cretaceous Period into the 21st Century.* Boston: Birkhäuser Verlag, 2001.
- Tyman, J. H. P. *Synthetic and Natural Phenols.* Amsterdam: Elsevier, 1996.
- Ware, George W. *The Pesticide Book.* Batavia, IL: Mesiter, 1999.
- Weinberg, Alan Bennet, and Bonnie K. Bealer. *The World of Caffeine: The Science and Culture of the World's Most Popular Drug.* New York: Routledge, 2002.
- Weissermel, Klaus, and Hans Jürgen Arpe. *Industrial Organic Chemistry.* Weinheim, Germany: Wiley VCH, 2003, 117–122.
- Wyman, Carolyn. *JELL-O: A Biography.* Fort Washington, PA: Harvest Books, 2001.

PERIODICALS

- “Another Old Fashioned Product Vindicates Itself.” *Medical Update* (October 1992): 6.
- Arnst, Catherine. “A Preemptive Strike against Cancer.” *Business Week* (June 7, 2004): 48.
- Baker, Linda. “The Hole in the Sky (Ozone Layer).” *E* (November 2000): 34.

- Bauman, Richard. "Getting Skunked: Understanding the Antics behind the Smell." *Backpacker* (May 1993): 30 31.
- Drake, Geoff. "The Lactate Shuttle Contrary to What You've Heard, Lactic Acid Is Your Friend." *Bicycling* (August 1992): 36.
- Fox, Berry. "Not Fade Away." *New Scientist* (March 1, 2003); 40.
- "Global Ethyl, Butyl Acetate Demand Expected to Rebound." *The Oil and Gas Journal* (April 24, 2000): 27.
- Gorman, Christine. "The Bomb Lurking in the Garden Shed." *Time* (May 1, 1995): 54.
- Karaim, Reed. "Not So Fast with the DDT: Rachel Carson's Warnings Still Apply." *American Scholar* (June 2005): 53 60.
- Keenan, Faith. "Blocking Liver Damage." *Business Week* (October 21, 2002): 147.
- Kluger, Jeffrey. "The Buzz on Caffeine." *Time* (December 20, 2004): 52.
- Lazear, N. R. "Polycarbonate: High Performance Resin." *Advanced Materials & Processes* (February 1995): 43 45.
- Legwold, Gary. "Hydration Breakthrough." *Bicycling* (July 1994): 72 73.
- Liu, Guanghua. "Chinese Cinnabar." *The Mineralogical Record* (January February 2005): 69 80.
- Malakoff, David. "Public Health: Aluminum Is Put on Trial as a Vaccine Booster." *Science* (May 26, 2000): 1323.
- Mardis, Anne L. "Current Knowledge of the Health Effects of Sugar Intake." *Family Economics and Nutrition Review* (Winter 2001): 88 91.
- McGinn, Anne Platt. "Malaria, Mosquitoes, and DDT: The Toxic War against a Global Disease." *World Watch* (May 1, 2002): 10 17.
- Metcalfe, Ed, et al. "Sweet Talking." *The Ecologist* (June 2000): 16.
- Milius, Susan. "Termites Use Mothballs in Their Nests." *Science News* (May 2, 1998): 228.
- Neff, Natalie. "No Laughing Matter." *Auto Week* (May 19, 2003): 30.
- "Nitroglycerin: Dynamite for the Heart." *Chemistry Review* (November 1999): 28.
- O'Neil, John. "And It Doesn't Taste Bad, Either." *New York Times* (November 30, 2004): F9.
- Pae, Peter. "Sobering Side of Laughing Gas." *Washington Post* (September 16, 1994): B1.
- Palvetz, Barry A. "A Bowl of Hope, Bucket of Hype?" *The Scientist* (April 2, 2001): 15.
- Rawls, Rebecca. "Nitroglycerin Explained." *Chemical & Engineering News* (June 10, 2002): 12.

For Further Information

- Rowley, Brian. "Fizz or Sizzle? Potassium Bicarbonate Could Help Spare Muscle and Bone." *Muscle & Fitness* (December 2002): 72.
- Russell, Justin. "Fuel of the Forgotten Deaths." *New Scientist* (February 6, 1993): 21 23.
- Schramm, Daniel. "The North American USP Petrolatum Industry." *Soap & Cosmetics*, (January 2002): 60 63.
- Stanley, Peter. "Nitric Oxide." *Biological Sciences Review* (April 2002): 18 20.
- "The Stink that Stays." *Popular Mechanics* (December 2004): 26.
- Strobel, Warren P. "Saddam's Lingering Atrocity." *U.S. News & World Report* (November 27, 2000): 52.
- "Strong Muscle and Bones." *Prevention* (June 1, 1995): 70 73.
- "Taking Supplements of the Antioxidant." *Consumer Reports* (September 2003): 49.
- Travis, J. "Cool Discovery: Menthol Triggers Cold Sensing Protein." *Science News* (February 16, 2002): 101 102.
- U.S. Department of Health and Human Services. "Methanol Toxicity." *American Family Physician* (January 1993): 163 171.
- "Unusual Thermal Defence by a Honeybee against Mass Attack by Hornets." *Nature* (September 28, 1995): 334 336.
- "USDA Approves Phosphate to Reduce Salmonella in Chicken." *Environmental Nutrition* (February 1993): 3.
- Vartan, Starre. "Pretty in Plastic: Pleather Is a Versatile, though Controversial, Alternative to eather." *E* (September October 2002): 53 54.
- "Vitamins: The Quest for Just the Right Amount." *Harvard Health Letter* (June 2004): 1.
- Walter, Patricia A. "Dental Hypersensitivity: A Review." *The Journal of Contemporary Dental Practice* (May 15, 2005): 107 117.
- Young, Jay A. "Copper (II) Sulfate Pentahydrate." *Journal of Chemical Education* (February 2002): 158.

WEBSITES

- "The A to Z of Materials." Sydney, Australia: Azom.com. <http://www.azom.com/> (accessed on March 1, 2006).
- "Air Toxics Website." U.S. Environmental Protection Agency Technology Transfer Network <http://www.epa.gov/ttn/atw/> (accessed on March 10, 2006).

Calorie Control Council. Atlanta, GA: Calorie Control Council.
<http://www.caloriecontrol.org/>

CHEC's HealthHouse: The Resource for Environmental Health Risks Affecting Your Children. Los Angeles, CA: Children's Health Environmental Coalition. <http://www.checnet.org/ehouse> (accessed on March 13, 2006).

Chemfinder.com. Cambridge, MA: CambridgeSoft Corporation.
<http://www.chemfinder.cambridgesoft.com> (accessed on March 13, 2006).

"Chemical Backgrounder." Itasca, IL: National Safety Council.
<http://www.nsc.org/library/chemical/> (accessed on March 1, 2006).

"Chemical Profiles." Scorecard. The Pollution Information Site. Washington, DC: Green Media Toolshed. http://www.scorecard.org/chemical_profiles/ (accessed on March 1, 2006).

The Chemistry Store. Cayce, SC: ChemistryStore.com, Inc. <http://www.chemistrystore.com/> (accessed on March 13, 2006).

Chemistry.org. Washington, DC: American Chemical Society.
<http://www.chemistry.org/portal/a/c/s/1/home.html> (accessed on March 13, 2006).

Cheresources.com: Online Chemical Engineering Information. Midlothian, VA: The Chemical Engineers' Resource Page.
<http://www.cheresources.com> (accessed on March 13, 2006).

"Dietary Supplement Fact Sheets." Bethesda, MD: U.S. National Institutes of Health, Office of Dietary Supplements.
http://ods.od.nih.gov/Health_Information/Information_About_Individual_Dietary_Supplements.aspx (accessed on March 13, 2006).

Drugs.com. Auckland, New Zealand: Drugsite Trust. <http://www.drugs.com> (accessed on March 13, 2006).

EnvironmentalChemistry.com. Portland, ME: Environmental Chemistry.com. <http://environmentalchemistry.com> (accessed on March 13, 2006).

Exploratorium: the museum of science, art and human perception. San Francisco, CA: Exploratorium at the Palace of Fine Arts.
<http://www.exploratorium.edu/> (accessed on March 13, 2006).

ExToxNet: The Extension Toxicology Network. Corvallis, OR: Oregon State University. <http://extoxnet.orst.edu/> (accessed on March 13, 2006).

Fibersource: The Manufactured Fibers Industry. Arlington, VA: Fiber Economics Bureau. <http://www.fibersource.com> (accessed on March 13, 2006).

For Further Information

- General Chemistry Online. Frostburg, MD: Frostburg State University, Department of Chemistry. <http://antoine.frostburg.edu/chem/senese/101/index.shtml> (accessed on March 13, 2006).
- Household Products Database. Bethesda, MD: U.S. National Library of Medicine. <http://householdproducts.nlm.nih.gov/> (accessed on March 13, 2006).
- Integrated Risk Information System. Washington, DC: U.S. Environmental Protection Agency. <http://www.epa.gov/iris/index.html> (accessed on March 13, 2006).
- "International Chemical Safety Cards (ICSCs)." Geneva, Switzerland: International Labour Organization. <http://www.ilo.org/public/english/protection/safework/cis/products/icsc/index.htm> (accessed on March 1, 2006).
- "International Chemical Safety Cards." Atlanta, GA: U.S. National Institute for Occupational Safety and Health. <http://www.cdc.gov/niosh/ipcs/icstart.html> (accessed on March 1, 2006).
- IPCS INTOX Databank. Geneva, Switzerland: International Programme on Chemical Safety. <http://www.intox.org/> (accessed on March 13, 2006).
- Kimball's Biology Pages. Andover, MA: John W. Kimball. <http://biologypages.info> (accessed on March 13, 2006).
- Linus Pauling Institute Micronutrient Information Center. Corvallis, OR: Oregon State University, Linus Pauling Institute. <http://lpi.oregonstate.edu/infocenter/> (accessed on March 13, 2006).
- MadSci Network. Boston, MA: Third Sector New England. <http://www.madsci.org> (accessed on March 13, 2006).
- Medline Plus. Bethesda, MD: U.S. National Library of Medicine. <http://www.nlm.nih.gov/medlineplus> (accessed on March 1, 2006).
- Mineral Information Institute. Golden, CO: Mineral Information Institute. <http://www.mii.org> (accessed March 13, 2006).
- Mineralogy Database. Spring, TX: Webmineral.com. <http://webmineral.com/> (accessed on March 13, 2006).
- "Molecule of the Month Page." Bristol, United Kingdom: University of Bristol School of Chemistry. <http://www.chm.bris.ac.uk/motm/motm.htm> (accessed on March 1, 2006).
- National Historic Chemical Landmarks. Washington, DC: American Chemical Society. <http://acswebcontent.acs.org/landmarks/index.html> (accessed on March 13, 2006).
- NIST Chemistry Webbook. Gaithersburg, MD: U.S. National Institute of Standards and Technology. <http://webbook.nist.gov> (accessed on March 13, 2006).

- PAN Pesticides Database. San Francisco, CA: Pesticide Action Network, North America. <http://www.pesticideinfo.org> (accessed March 13, 2006).
- “Polymer Science Learning Center.” Hattiesburg, Mississippi: University of Southern Mississippi, Department of Polymer Science. <http://www.psclc.ws/> (accessed on March 13, 2006).
- PubChem. Bethesda, MD: U.S. National Library of Medicine. <http://pubchem.ncbi.nlm.nih.gov/> (accessed on March 13, 2006).
- Reciprocal.net. Bloomington, IN: Indiana University Molecular Structure Center. <http://www.reciprocalnet.org/> (accessed on March 13, 2006).
- The Science Center: A Teacher’s Guide to Educational Resources on the Internet. Arlington, VA: Chlorine Chemistry Council. <http://www.scienceeducation.org> (accessed on March 13, 2006).
- Shakhashiri, Bassam Z. “Chemical of the Week.” Science Is Fun. Madison, WI: University of Wisconsin. <http://scifun.chem.wisc.edu/chemweek/chemweek.html> (accessed on March 13, 2006).
- 3Dchem.com. Oxford, United Kingdom: University of Oxford, Department of Chemistry. <http://www.3dchem.com> (accessed on March 13, 2006).
- “ToxFAQs™” Atlanta, GA: U.S. Agency for Toxic Substances and Disease Registry. <http://www.atsdr.cdc.gov/toxfaq.html> (accessed on March 13, 2006).
- “Toxicological Profile Information Sheets.” Atlanta, GA: U.S. Agency for Toxic Substances and Disease Registry. <http://www.atsdr.cdc.gov/toxpro2.html> (accessed on March 1, 2006).
- TOXNET: Toxicology Data Network. Bethesda, MD: U.S. National Library of Medicine. <http://toxnet.nlm.nih.gov> (accessed on March 13, 2006).
- World of Molecules Home Page. <http://www.worldofmolecules.com/> (accessed on March 13, 2006).

index

This index is sorted word by word. Italic type indicates volume numbers; boldface indicates main entries; (ill.) indicates illustrations.

A

- Abbott Laboratories, 3:742
Abrasives, 1:50
Acetaminophen, 1:19-22, 20 (ill.)
Acetic acid, 1:23-26, 24 (ill.)
Acetone, 1:289-291
Acetylene, 1:27-30, 28 (ill.)
Acetylsalicylic acid, 1:31-35, 32 (ill.)
Acid rain, 1:149, 153; 2:499; 3:822
Activated charcoal, 2:457
Adenosine, 1:140
Advil, 1:11
Alarm pheromones, 1:127
Albrecht, H.O., 2:407
Alcohol, 2:297-301, 449-453
Aleve, 2:477
Alka-Seltzer, 3:726
Alkenes, 2:307-311
Alpha-tocopherol, 1:37-40, 38 (ill.)
Aluminum fluoride, 1:41-44, 42 (ill.)
Aluminum hydroxide, 1:45-48, 46 (ill.), 50
Aluminum metal, 1:42, 50-51
Aluminum oxide, 1:49-52, 50 (ill.)
Aluminum potassium sulfate, 1:53-56, 54 (ill.)

Index

- Amatol, 1:18
Ammonia, 1:57-61, 58 (ill.), 69-71, 73-74; 2:446; 3:761-762, 868-869
Ammonia-soda process, 1:64
Ammonium chloride, 1:63-67, 64 (ill.)
Ammonium hydroxide, 1:69-71, 70 (ill.)
Ammonium nitrate, 1:73-76, 74 (ill.); 2:495
Ammonium perchlorate, 2:541-542
Ammonium sulfate, 1:77-80, 78 (ill.)
Amoxicillin, 1:81-84, 82 (ill.)
Amyl acetate, 1:85-88, 86 (ill.)
Amyl nitrite, 1:89-92, 90 (ill.)
Analgesics, 1:9-13, 19-22, 31-34
Andrews, Edmund, 2:514
Anesthetics, 1:211-212; 2:514-515
ANFO (Ammonium nitrate fuel oil), 1:75
Angina, treatment of, 1:90
Anhydrous calcium sulfate, 1:165-169
Animal feeds, 3:652, 661, 870
Antacids
 aluminum hydroxide, 1:45-46
 calcium carbonate, 1:144
 magnesium hydroxide, 2:416-417
 potassium bicarbonate, 3:621-622
 sodium bicarbonate, 3:726
Anti-inflammatories, 1:9
Antibiotics, 1:81-84; 2:535-539
Antifouling agents, 1:248
Antifreeze, 2:315
Antioxidants, 1:38, 110, 135
Antiseptics, 1:117, 261; 2:365; 3:652, 859-865
Arm & Hammer, 3:724
Armstrong, Henry E., 1:1
Aromatic hydrocarbons, 1:255-258; 2:303-306
Ascorbic acid, 1:93-97, 94 (ill.)
Aspartame, 2:401-405
Asphyxiants, 1:27-30
Aspirin, 1:31-35
Astringents, 1:54; 3:834
Atlantic Richfield Corp., 2:459
Avery, Charles E., 2:392
Azeotropic mixtures, 2:459-461
- Bad breath, 2:457
Baekeland, Leo Hendrik, 2:561
Baking soda, 3:723-727
Banks, Robert, 2:580, 582, 588
Baratol, 1:18
Barnes, W., 1:271
BASF Corp., 1:28; 3:598
Bauxite, 1:45, 49
Bayer, Otto, 3:610
Bayer AG Co., 1:33; 2:576
Beadle, Clayton, 1:196, 207
Beecham lab, 1:83
Bees, 2:379
Béguin, Jean, 1:289
Ben-Gay, 1:173
Benzene, 1:99-103, 100 (ill.); 3:855
Benzoic acid, 1:105-108, 106 (ill.)
Beriberi, 3:847-851
Berthollet, Claude Louis, 1:57; 3:760
Berzelius, Jons Jakob, 3:853
Beta-carotene, 1:109-113, 110 (ill.); 3:678-680
Bevan, Edward John, 1:196, 207
B.F. Goodrich, 2:384; 3:616-617
BHA, 1:133-136
BHT, 1:133-136
Bishop, Katherine Scott, 1:38
Bitrex, 1:271-274
Black, Joseph, 1:151, 177-178
Black Leaf 40, 2:489
Black mercury(II) sulfide, 2:439-442
Bleaches, 2:365; 3:761, 796, 820
Block copolymers, 2:571-574
Blyth, Alexander Wynter, 3:683
Bock, Walter, 2:572
Bondt, N., 2:307
Boots Co., 1:9-11
Borax, 1:116; 3:789-792
Boric acid, 1:115-118, 116 (ill.)
Bosch, Carl, 1:58
Boyle, Robert, 2:451, 513; 3:659
Braconnot, Henri, 1:196
Breast implants, 2:594
Bristol-Meyers, 1:11
Brown-Séquard, 3:837-838

B

Brucite, 2:415
 Brunton, Sir Thomas Lauder, 2:508
 Büchner, Johann, 1:32
 Buffering agents, 3:721
 Building products, chemicals in
 calcium carbonate, 1:145
 calcium hydroxide, 1:148
 calcium silicate, 1:162-163, 167-168
 polystyrene, 3:599
 polyvinyl chloride, 3:618
 silicon dioxide, 3:698
 Bussy, Antoine, 1:289
 Butane, 1:119-123, 120 (ill.)
 Butenandt, Adolf Friedrich Johann, 3:838
 Butlerov, Alexander Mikhailovich, 2:325
 Butyl acetate, 1:125-128, 126 (ill.)
 Butyl mercaptan, 1:129-132, 130 (ill.)
 Butylated hydroxyanisole (BHA), 1:133-136,
 134 (ill.)
 Butylated hydroxytoluenene (BHT), 1:133-136,
 134 (ill.)

C

Caffeine, 1:137-141, 138 (ill.); 3:843-846
 Calcium carbonate, 1:143-146, 144 (ill.), 151-152,
 178
 Calcium hydroxide, 1:147-150, 148 (ill.)
 Calcium magnesium acetate, 1:25
 Calcium oxide, 1:151-155, 152 (ill.)
 Calcium phosphate, 1:155-159, 156 (ill.)
 Calcium silicate, 1:161-164, 162 (ill.)
 Calcium sulfate, 1:165-169, 166 (ill.)
 Calcium sulfate dihydrate, 1:165-169
 Calcium sulfate hemihydrate, 1:165-169
 Campho-phenique, 1:173
 Camphor, 1:171-175, 172 (ill.)
 Carbohydrates. *See* Sugars
 Carbolic acid, 2:559-563
 Carbon dioxide, 1:177-181, 178 (ill.)
 Carbon monoxide, 1:183-187, 184 (ill.)
 Carbon tetrachloride, 1:189-193, 190 (ill.)
 Carbonation of beverages, 1:179
 Carcinogens
 benzene, 1:102
 carbon tetrachloride, 1:192

chloroform, 1:214
 dichlorodiphenyltrichloroethane (DDT),
 1:283-287
 ethylene oxide, 2:320
 formaldehyde, 2:327
 gamma-1,2,3,4,5,6-hexachlorocyclohexane,
 2:335
 isoprene, 2:384
 methyl-t-butyl-ether, 2:462
 polyvinyl chloride, 3:619
 saccharin, 3:691
 styrene, 3:805
 triclocarban, 3:861
 triclosan, 3:865
 trinitrotoluene, 1:18
 Carotenoids, 1:109-113
 Carothers, Wallace, 2:519
 Carson, Rachel, 1:284
 Cartier, Jacques, 1:95
 Cavendish, Henry, 2:494
 Caumont, Joseph Bienaimé, 1:137, 219
 Celluloid, 1:202-203
 Cellulose, 1:195-199, 196 (ill.)
 Cellulose nitrate, 1:201-205, 202 (ill.)
 Cellulose xanthate, 1:207-210
 Cement, 1:145, 162-163, 167-168
 Chain, Ernst, 2:536
 Chalk, 1:143
 Chemical intermediaries, 1:119, 121-122, 131
 Chemical manufacturing, chemicals used in
 acetylene, 1:28
 benzene, 1:101
 carbon monoxide, 1:185
 cumene, 1:255-256
 ethylene, 2:309
 ethylene oxide, 2:317-318
 hydrogen chloride, 2:359
 methane, 2:446
 methyl alcohol, 2:451
 methyl mercaptan, 2:456
 propylene, 3:671
 sodium chloride, 3:737
 sulfur dioxide, 3:820
 toluene, 3:855
 Chemical warfare, 1:5-7
 Chemiluminescence, 2:407
 Chevreul, Eugène, 1:223; 2:349

Index

- Chlorine, 2:358; 3:795, 3:786
Chlorofluorocarbons (CFCs), 1:190-191, 277-281
Chloroform, 1:211-215, 212 (ill.)
Chlorophyll, 1:217-221, 218 (ill.)
Chocolate, 1:138; 3:843-846
Cholesterol, 1:223-227, 224 (ill.)
Church, Austin, 3:724
Cinnabar, 2:439
Cinnamaldehyde, 1:229-232, 240 (ill.)
Cinnamomum camphora, 1:171, 173
Cinnamomum zeylanicum, 1:229
Citric acid, 1:233-237, 234 (ill.)
Cleaning products, chemicals in
 ammonium hydroxide, 1:70
 denatonium benzoate, 1:274
 oxalic acid, 2:527
 sodium bicarbonate, 3:726
 sodium hydroxide, 3:755
 sodium hypochlorite, 3:761
 sodium perborate, 3:766-767
 sodium silicate, 3:780
 tribasic sodium phosphate, 3:771
 See also Soaps and detergents, chemicals in
Clinton Corn Processing Co., 2:330-331
Clorox, 3:761
Cloud seeding, 3:702-704
Coal tar, 2:473-475
Cocoa beans, 3:843-846
Coffee, 1:137-141
Cola drinks, 1:137-141
Collagen, 1:239-242; 2:338
Contact process, 3:827
Coolants, 2:315
Coover, Harry, 1:259-260
Copper(I) oxide, 1:243-246, 244 (ill.)
Copper(II) oxide, 1:247-250, 248 (ill.)
Copper(II) sulfate, 1:251-254, 252 (ill.)
Cotton, 1:195-198
Cox enzymes, 1:12; 2:479-480
Cracking process, 2:309, 382; 3:669-670
Cream of tartar, 3:629-631
Crede, Carl, 3:705
Crest toothpaste, 3:800
Cross, Charles Frederick, 1:196, 207
Cross-Bevan-Beadle method, 1:208
Crude oil, 2:548, 553-557
Cruikshank, William, 1:184
Cumene, 1:101, 255-258, 256 (ill.)
Currie, James, 1:235
Cyanamide process, 1:58
Cyanoacrylate, 1:259-263, 260 (ill.)
Cyanocobalamin, 1:265-269, 266 (ill.)
Cyclamates, 3:741-744
Cyclohexane, 1:101

D

- Daguerre, Louis Jacques Mandé, 3:702
Darby, William J., 2:322
Davy, Edmund, 1:27
Davy, Sir Humphry, 2:358, 513, 515
Day, Harry G., 3:800
DDT, 1:283-287
de la Salle, Pouletier, 1:223
de Lassone, Joseph Marie François, 1:184
Dead-burning, 2:419-420
Death penalty, 3:641
Decaffeination, 1:139
DEET, 2:469-471
Deicers, 1:25; 3:721
Deiman, J.R., 2:307
Denatonium benzoate, 1:271-275, 272 (ill.)
Dentistry, chemicals used in
 cyanoacrylate, 1:261
 dibasic calcium phosphate, 1:157
 nitrous oxide, 2:514-515
 sodium fluoride, 3:748-749
 stannous fluoride, 3:799-802
Deodorizers, room, 1:90
Diabetes, 2:346
Diapers, disposable, 3:773-776
Diazonium compounds, production of, 1:91
Dibasic calcium phosphate, 1:155-159
Dibasic sodium phosphate, 3:769-772
Dichlorodifluoromethane, 1:277-281, 278 (ill.)
Dichlorodiphenyltrichloroethane (DDT),
 1:283-287, 284 (ill.)
Diesel fuel, 2:555
Dimethyl ketone, 1:289-291, 290 (ill.)
Dow, Herbert Henry, 2:412
Dow Chemical Co., 2:584; 3:598
Dow process, 2:412
Drain cleaners, 3:755

Dry cell batteries, 1:65
 Drying agents, 3:698
 Dumas, Jean Baptiste André, 1:211, 229, 289
 DuPont Chemical Corp., 1:277-278; 2:519; 3:603-604, 742
 Dwight, John, 3:724
 Dyes and pigments
 2,4,6-trinitrotoluene in, 1:18
 aluminum potassium sulfate, 1:55
 copper(I) oxide, 1:244
 copper(II) oxide, 1:247-248
 iron(II) oxide, 2:367
 iron(III) oxide, 2:373
 mercury(II) oxide, 2:441
 See also Paint, chemicals in
 Dynamite, 2:508

E

Eastman Kodak, 1:259-260
 Eggs, preservation of, 3:781
 Ehlers-Danlos syndrome, 1:240
 Eijkman, Christiaan, 3:849
 Elvehjem, Conrad Arnold, 2:484
 Engel-Precht process, 3:634-635
 Environmental concerns
 acid rain, 1:153, 149
 chlorofluorocarbons (CFCs), 1:190-191, 277-281
 dichlorodiphenyltrichloroethane (DDT), 1:283-287
 methyl-t-butyl ether, 2:461
 nitric oxide, 2:497-499
 nitrogen dioxide, 2:504-505
 ozone layer depletion, 1:190-191, 213-214, 279
 petroleum, 2:556
 phosphates, 2:568
 pollution control, 1:152-153
 polystyrene, 3:600
 polyvinyl chloride, 3:618-619
 sulfur dioxide, 3:822
 triclocarban, 3:862
 triclosan, 3:865
 Epsom salts, 2:429-433
 Erlenmeyer, Emil, 1:229; 2:473

Esters, 1:259-263
 Ethane, 1:120-121
 Ethyl acetate, 1:293-295, 294 (ill.)
 Ethyl alcohol, 2:297-301, 298 (ill.)
 Ethylbenzene, 1:101; 2:303-306, 304 (ill.)
 Ethylene, 1:2; 2:307-311, 308 (ill.), 581-582; 3:666
 Ethylene glycol, 2:313-316, 314 (ill.), 318
 Ethylene oxide, 2:317-320, 318 (ill.)
 Eutrophication, 2:568
 Evans, Herbert McLean, 1:37-38
 Exhaust gases, treatment of, 1:148
 Explosives
 acetylene, 1:29
 ammonium nitrate, 1:74-75
 cellulose nitrate, 1:204
 nitric acid, 1:60
 nitroglycerin, 2:507-510
 perchlorates, 2:542-543
 potassium nitrate, 3:656
 styrene, 3:805
 TNT, 1:15-18
 toluene, 3:855-856

F

Fabrics
 bleaching, 3:760
 cotton, 1:195-198
 dyeing of, 1:55; 3:720
 flax, 1:195-198
 nylon, 2:519-523
 polypropylene, 2:589
 polyurethane fibers, 3:612
 rayon, 1:207-210
 stain and water resistance, 3:605-606
 Fahlberg, Constantine, 3:689-690
 Faraday, Michael, 1:99
 Farben, I.G., 3:598
 Fats, artificial, 3:813-816
 Fawcett, Eric, 2:580
 FDA (Food and Drug Administration, U.S.), 1:260-261; 2:404; 2:404; 3:815
 Fermentation, 1:294; 2:298-301
 Fertilizer, chemicals in
 ammonia, 1:59-60

Index

- ammonium nitrate, 1:74; 2:495
ammonium sulfate, 1:77-78
monocalcium phosphate, 1:157
phosphoric acid, 2:567
potassium chloride, 3:641
potassium sulfate, 3:660
sulfuric acid, 3:827
tricalcium phosphate, 1:158
urea, 3:870
- Film, photographic, 1:203-204
Fingerprints, collection of, 1:262
Fire extinguishers, 1:180; 3:622
Firestone, 2:384
Fireworks, 1:75; 2:542-543; 3:656
Fischer, Emil, 1:138; 2:345
Fischer, Hans, 1:219
Fischer-Tropsch reactions, 1:185
Fittig, Rudolf, 3:855
Flammable Fabrics Act of 1953, 1:209
Flavorings
 amyl acetate, 1:85-86
 cinnamaldehyde, 1:231
 ethyl acetate, 1:293-295
 isoamyl acetate, 2:377-380
 salt and salt substitutes, 2:465-468; 3:641, 653, 735-739
 sugars, 2:329-332, 343-347; 3:807-811
 sweeteners, artificial, 2:349-352, 401-405; 3:689-693, 741-744
 vanillin, 3:873-876
- Flax, 1:195-198
Fleming, Alexander, 1:82; 2:535
Florey, Howard, 2:536
Flourens, Marie Jean-Pierre, 1:211
Fluoridating agents, 3:644
Fluoridation, water, 1:43
Fluorides, dental, 3:747-751, 799-802
Foams, 3:611-612
Folate, 2:321-324
Folic acid, 2:321-324, 322 (ill.)
Folkers, Karl, 1:266-267; 3:674
Food additives
 acidification agents, 1:158; 3:641, 771
 buffering agents, 3:721
 colorings, 1:112
 cream of tartar, 3:630-631
 folic acid, 2:321-324
- isoamyl acetate, 2:377-380
jelling agents, 2:531-533
lactic acid, 2:393
lactose, 2:397-399
leavening agents, 3:723, 726
nutritional supplements, 1:158
potassium bicarbonate, 3:621-622
potassium bisulfate, 3:625-627
preservatives, 1:107, 117, 133-135, 233-236; 3:786-787
stimulants, 1:117, 137-141
 See also Flavorings; Vitamins
- Food and Drug Administration, U.S. (FDA), 1:260-261; 2:404; 3:815
- Food packaging, 3:775
Food production, 3:649
Forensic science, 2:409
Formaldehyde, 2:325-328, 326 (ill.)
Fox, Daniel W., 2:577
Free radicals, 1:38
Freon, 1:214, 277-281
Friedel, Charles, 2:349
Frohlich, Theodore, 1:95
Fructose, 2:329-332, 330 (ill.)
Fruits, 2:531-532
Fuels
 additives, 2:450, 459-462, 515-516
 butane, 1:121
 ethyl alcohol, 2:299
 gasoline, 2:555
 industrial, 1:185
 methane, 2:445
 oil, 1:75; 2:555
 perchlorates, 2:542-543
 propane, 3:663-667
- Fumigants, 1:252
Fungicides, 1:117, 244
Funk, Casimir, 2:483
Fusion, 2:420-421
- G**
- Gamma-1,2,3,4,5,6-hexachlorocyclohexane, 2:333-336, 334 (ill.)
Gasoline, 2:555
Gatorade, 3:744

Gay-Lussac, Louis-Joseph, 1:116
 G.D. Searle Co., 2:401-402, 2:404
 Geber. *See* Jabir ibn Hayyan (Geber)
 Gelatin, 2:337-341
 Gemstones, 1:50
 Gerhardt, Charles Frederic, 1:19, 32; 2:559
 Gibson, Reginald, 2:580
 Glacial acetic acid, 1:23-26
 Glass manufacturing, chemicals used in
 boric acid, 1:117
 potassium carbonate, 3:636
 potassium fluoride, 3:645
 silicon dioxide, 3:698
 silver oxide, 3:711-712
 sodium carbonate, 3:731-732
 sodium tetraborate, 3:791
 Glass substitutes, 2:584
 Glauber, Johann, 1:73, 75; 2:358, 493, 559; 3:659
 Glazes, 1:244; 3:716
 Glow sticks, 2:410
 Glucose, 2:343-347, 344 (ill.)
 Glycerol, 2:349-352, 350 (ill.)
 Goldeberger, Joseph, 2:484
 Gout, 2:394-395
 Green chemistry, 1:11
 Grignard, Victor, 2:592
 Grijins, Gerrit, 3:849
 Gum benzoin, 1:105
 Guncotton, 1:201-205
 Gunpowder, 3:656-657
 Guthrie, Frederick, 1:5
 Guthrie, Samuel, 1:211
 Gypsum, 1:165-168

H

Haber, Fritz, 1:58-59
 Haber-Bosch process, 1:58-59, 178
 Hall, Charles Martin, 1:50-51
 Hallwachs, Wilhelm, 1:245
 Hard-burning, 2:420-421
 Harris, S.A., 3:674
 Haworth, Sir Walter Norman, 1:95
 HDPE (high density polyethylene), 2:579, 2:582
 Heart attack, prevention and treatment,
 1:33-34; 2:500, 510

Helmont, Jan Baptista van, 1:177; 2:497;
 3:825-826
 Hematite, 2:371-373
 Hexachlorophene, 3:860
 Hexane, 2:353-356, 354 (ill.)
 High-density lipoproteins (HDL), 1:226
 High-fructose corn syrup, 2:330-331
 Hodgkin, Dorothy, 1:267
 Hoffman, Felix, 1:33
 Hofmann, August Wilhelm von, 2:325; 3:853
 Hogan, John, 2:580, 582, 588
 Holberg, Wilhelm, 1:116
 Holst, Alex, 1:95
 Hormones, plant, 2:307-308
 Hot springs, 1:115-118; 2:430-431
 Hula hoops, 2:581
 Human body
 cholesterol, 1:225-227
 collagen, 1:239-241
 fetal development, 2:321-323
 hormones, 3:837-840
 lactic acid, 2:392-393
 nicotine, effects of, 2:489-490
 nitric oxide in, 2:499-500
 nitroglycerin, use of, 2:510
 salt and, 3:738-739
 sucrose, effects of, 3:811
 urea in, 3:867-870
 water in, 3:882
 See also Medicine and medical products;
 Vitamins
 Hyatt, John Wesley, 1:202-203
 Hydrocarbons, 2:553-554
 Hydrochloric acid, 2:357-361
 Hydrochlorofluorobarbons (HCFCs), 1:212-214
 Hydrogen chloride, 2:357-361, 358 (ill.)
 Hydrogen peroxide, 2:363-366, 364 (ill.); 3:765
 Hydroxyapatite, 1:157

I

Ideda, Kikunae, 2:465
 Imperial Chemical Industries (ICI), 2:580
 Industrial accidents, 1:75
 Insect repellants, 2:469-471
 Insecticides, 1:283-287; 2:335, 488-489

Index

Iron(II) oxide, 2:367-369, 368 (ill.)
Iron(III) oxide, 2:371-375, 372 (ill.)
Isler, Otto, 3:679
Isoamyl acetate, 2:377-380, 378 (ill.)
Isoprene, 2:381-385, 382 (ill.)
Isopropyl alcohol, 2:387-389, 388 (ill.)

J

Jabir ibn Hayyan (Geber), 1:23, 234-235
Jams and jellies, 2:532-533
JELL-O, 2:337-338
Joiner, Fred, 1:259-260

K

Kaminsky, Walter, 2:581-582
Karrer, Paul, 1:38, 110; 3:678, 683
Kekulé, Friedrich August, 1:99-100, 102; 2:326
Kerosene, 2:555-556
Kidd, John, 2:473
Kidney stones, 1:157
King, Charles Glen, 1:95
Kipping, Frederic Stanley, 2:592
Kirchhof, Gottlieb Sigismund Constantine, 2:343
Klate, Friedrich Heinrich August, 3:616
Knox Gelatin, 2:337-338
Krazy Glue, 1:259-263
Krebs, Sir Hans Adolf, 1:234
Krumbhaar, E.B., 1:7
Krumbhaar, H.D., 1:7
Kuhn, Richard, 3:674, 679, 683

L

L-aspartyl-L-phenylalanine methyl ester, 2:401-405, 402 (ill.)
Lactic acid, 2:391-395, 392 (ill.)
Lactose, 2:397-400, 398 (ill.)
Laughing gas, 2:513-517
Lauwerenburgh, A., 2:307
LDPE (low density polyethylene), 2:579, 582

Lead chamber process, 3:826
Leavening agents, 3:723, 726
Lebedev, Sergei, 1:1-2
Leblanc, Nicolas, 2:358; 3:730-731
Leroux, Henri, 1:32
Lethal injections, 3:641
Levenstein process, 1:6
Libavius, Andreas, 1:79
Liebig, Justus von, 1:289
Light-burning, 2:420-421
Lignin, 3:874-876
Lime, 1:151-155
Limestone, 1:144
Lind, James, 1:95
Liniments, 1:173
Liquid Bandage, 1:261
Liquid paraffin, 2:548
Liquid petroleum gas (LPG), 1:121
LoSalt, 3:641

Low-density lipoproteins (LDL), 1:226
Lucite, 2:584
Luminol, 2:407-410, 408 (ill.)
Lye, 3:753-757

M

Macfarlan Smith Corp., 1:272
Magnesium chloride, 2:411-414, 412 (ill.)
Magnesium hydroxide, 2:415-417, 416 (ill.)
Magnesium metal, production of, 2:412
Magnesium oxide, 2:419-422, 420 (ill.)
Magnesium silicate hydroxide, 2:423-427, 424 (ill.)
Magnesium sulfate, 2:429-433, 430 (ill.)
Malaria, 1:283-286
Marggraf, Andreas Sigismund, 2:343
Marlex, 2:588
Mattson, Fred, 3:813
Maumee Chemical Co., 3:691
Maynard, J. P., 1:204
McKay, Frederick S., 3:801
Medicine and medical products
 analgesics, 1:9-13, 19-22, 31-34, 173
 antacids, 1:45-46, 144; 2:416-417;
 3:621-622, 726
 antibiotics, 1:81-84; 2:535-538

antiseptics, 1:117, 173, 261; 2:365; 3:652, 859-865
aspirin, 1:31-34
astringents, 3:834
caffeine, 1:137-141
cough and sinus treatments, 2:437, 562
Epsom salts, 2:431
hormones and hormone treatments, 2:543; 3:837-840
laxatives, 2:550
liniments, 1:173
nitroglycerin, 2:500, 510
NSAIDs, 2:477-480
polyvinyl chloride in, 3:618
silver nitrate, 3:707-708
skin treatments, 2:437-438, 548-549, 562
SSKI, 3:652
theobromine, 3:845-846
vasodilators, 1:90
wound sealants, 1:260-261
See also Dentistry, chemicals used in
Melsens, Louise, 2:487
Ménard, Louis-Phillippe, 1:204
Menthol, 2:435-438, 436 (ill.)
Mercury metal, 2:441
Mercury(II) sulfide, 2:439-442, 440 (ill.)
Metals, refining, 1:186
Methane, 1:120-121; 2:443-447, 444 (ill.)
Methanol, 1:272; 2:449-453
Methyl alcohol, 2:449-453, 450 (ill.)
Methyl mercaptan, 2:455-458, 456 (ill.)
Methyl-t-butyl ether, 2:450, 459-463, 460 (ill.)
Methylxanthines, 1:138
MFP (Sodium monofluorophosphate), 3:801
Midgley, Thomas, 1:278, 280
Military dynamite, 1:17
Milk, 2:397-399
Milk of Magnesia, 2:415-417
Miller, A. K., 1:1
Mineral oil, 2:548
Minerals (nutritional), 1:143-146, 155-159
Minot, George, 1:266
Mitchell, Henry K., 2:322
Mitscherlich, Eilhardt, 1:99
Mohammad ibn Zakariya al-Razi, 3:825
Mohs, Frederick, 1:51
Molina, Mario, 1:190, 278

Monobasic calcium phosphate, 1:155-159
Monobasic sodium phosphate, 3:769-772
Monosodium glutamate (MSG), 2:465-468, 466 (ill.)
Monsanto Corp., 1:24
Montecatini, 2:588
Montreal Protocol on Substances that Deplete the Ozone Layer, 1:191, 214, 279
Morse, Harmon Northrop, 1:19
Mortar, 1:148
Mothballs, 2:475
Motrin, 1:9-10
MTBE (methyl-tert-butyl ether), 2:450, 459-463
Mueller, Joseph C., 3:799-800
Müller, Paul Hermann, 1:283
Mummification, 3:730
Murphy, William, 1:266
Murray, Sir James, 2:415
Muscles, 2:392-393
Muspratt, James, 3:853
Mustard gas, 1:5-8

N

n-butyl acetate, 1:125-126
n-butyl mercaptan, 1:129-132
Nail polish remover, 1:290
Napalm, 2:475
Naphthalene, 2:473-476, 474 (ill.)
Naproxen, 2:477-481, 478 (ill.)
Nascent oxygen, 2:365
Natta, Giulio, 2:384, 588-589
Natural gas, 1:119-120; 2:443-446; 3:663-665
Neljubow, Dmitri N., 2:308
Neutralizing agents, 1:148
Niacin, 2:483-486, 484 (ill.)
Nicot, Jean, 2:487
Nicotine, 2:487-491, 488 (ill.)
Nitrating agents, 2:505
Nitric acid, 1:60, 73-74; 2:493-496, 494 (ill.)
Nitric oxide, 2:497-501, 498 (ill.), 504
Nitro, 2:515-516
Nitrogen, 1:74, 78
Nitrogen dioxide, 2:503-506, 504 (ill.)
Nitroglycerin, 2:500, 507-511, 508 (ill.)
Nitrous oxide, 2:513-517, 514 (ill.)

Index

N,N-diethyl-3-methylbenzamide (DEET),
2:469-472, 470 (ill.)
Nobel, Alfred, 2:508-509
NSAIDs (nonsteroidal anti-inflammatory
drugs), 1:11-12, 19-22; 2:477-481
Nuisance dust, 1:50-51
Nuprin, 1:11
NutraSweet, 2:401-405
Nylon 6 and nylon 66, 2:519-523, 520 (ill.)

O

Obesity, 3:811
Odors and odorants, 1:130-131, 173; 2:377-380,
455-456
Oil of vitriol, 3:825-829
Oklahoma City bombing, 1:75
Olestra, 3:813-816
1,3-Butadiene, 1:1-4, 2 (ill.)
Ostwald, Wilhelm, 1:60; 2:494
Oxalic acid, 2:525-529, 526 (ill.)
Oxidizing agents, 2:505
Ozone layer, 1:190-191, 279

P

Paint, chemicals in, 1:145, 244; 3:886
See also Dyes and pigments
Paint thinners, 1:256
Paper manufacturing, chemicals used in
aluminum potassium sulfate, 1:55
calcium carbonate, 1:145
cellulose, 1:198
hydrogen peroxide, 2:365
sodium hydroxide, 3:755
sodium silicate, 3:781
sodium sulfite, 3:785-786
sodium thiosulfate, 3:795
Parkes, Alexander, 1:202
Pauling, Linus, 1:96
Payen, Anselme, 1:196
Pearl ash, 3:633-637
Pechmann, Hans von, 2:580
Pectin, 2:531-534
Péligot, Eugène Melchior, 1:229

Pellagra, 2:483-485
Pelletier, Pierre Joseph, 1:137-138, 219; 3:853
Pelouze, Théophile-Jules, 2:349
Penicillin, 1:81-83; 2:535-539, 536 (ill.)
Pennies, 1:249
Pentolite, 1:17
Peppermint, 2:435-436
Perchlorates, 2:541-545, 542 (ill.)
Periclase, 2:419
Pernicious anemia, 1:265-268
Perrin, M.W., 2:580
PET (polyethylene terephthalate), 2:315
Petrolatum, 2:547-551
Petroleum, 2:353-354, 473-475, 553-557;
3:663-665, 671
Petroleum jelly, 2:547-551
PFOA (perfluorooctanoic acid), 3:606-607
Phenol, 2:559-563, 560 (ill.)
Phenylketonuria, 2:404
Phillips, Charles Henry, 2:415
Phillips, Peregrine, 3:827
Phillips Petroleum, 2:580, 588
Phosgene, 1:191
Phosphates, 2:568
Phosphoric acid, 2:565-569, 566 (ill.); 3:827
Photoelectric cells, 1:244-245
Photography, 3:702, 706-707, 796
Photosynthesis, 1:217-221; 2:345
Pictet, A., 2:487
Pigments. *See* Dyes and pigments
Piria, Raffaele, 1:32
Plants, 1:195-199, 217-221; 2:307-308, 310;
3:831, 833
See also Photosynthesis
Plaster of Paris, 1:165-168
Plastic surgery, 1:240
Plastics manufacturing, chemicals used in,
1:107
benzoic acid, 1:107
ethylene, 2:309
ethylene glycol, 2:325-327
formaldehyde, 2:325-327
nylon, 2:521
1,3-butadiene, 1:3
petroleum, 2:556
poly(styrene-butadiene-styrene), 2:572-573
sodium hydroxide, 3:755

styrene, 3:804-805
 Plexiglas, 2:584, 2:585
 Plunkett, Roy J., 3:604
 PMMA (Polymethylmethacrylate), 2:583-586
 Poisoning, treatment of, 1:90
 Pollution. *See* Environmental concerns
 Polycarbonates, 2:575-578, 576 (ill.)
 Polyethylene, 2:579-582, 580 (ill.)
 Polymethylmethacrylate (PMMA), 2:583-586,
 584 (ill.)
 Polypropylene, 2:587-590, 588 (ill.); 3:669, 671
 Polysiloxane, 2:591-596, 592 (ill.)
 Polystyrene, 3:597-601, 598 (ill.), 804
 Poly(styrene-butadiene-styrene), 2:571-574,
 572 (ill.)
 Polytetrafluoroethylene, 3:603-607, 604 (ill.)
 Polyurethane, 3:609-613, 610 (ill.)
 Polyvinyl acetates, 1:107
 Polyvinyl chloride, 3:615-619, 616 (ill.)
 Poppers, 1:91
 Portland cement, 1:162-163, 167-168
 Potash, 3:633-637
 Potash, caustic, 3:647-650
 Potassium benzoate, 1:107
 Potassium bicarbonate, 3:621-623, 622 (ill.)
 Potassium bisulfate, 3:625-627, 626 (ill.)
 Potassium bitartrate, 3:629-631, 630 (ill.)
 Potassium carbonate, 3:633-637, 634 (ill.)
 Potassium chloride, 3:639-642, 640 (ill.)
 Potassium compounds, production of, 3:648
 Potassium fluoride, 3:643-645, 644 (ill.)
 Potassium hydroxide, 3:647-650, 648 (ill.)
 Potassium iodide, 3:651-654, 652 (ill.)
 Potassium nitrate, 3:655-658, 656 (ill.)
 Potassium perchlorate, 2:541-543
 Potassium sulfate, 3:659-662, 660 (ill.)
 Preservatives, food
 benzoic acid, role in, 1:107
 BHA and BHT, 1:134-135
 boric acid, 1:117
 citric acid, 1:233-236
 sodium sulfite, 3:786-787
 Priestley, Joseph, 1:57, 178, 184; 2:358, 513;
 3:821
 Procter & Gamble, 3:813-815
 Propane, 1:120-121; 3:663-667, 664 (ill.)
 Propellants, 1:122

Propylene, 3:666, 669-672, 670 (ill.)
 Provitamins, 1:110
 PVAs, 1:107
 PVC (polyvinyl chloride), 3:615-619
 Pyrex glass, 3:791
 Pyridoxine, 3:673-676, 674 (ill.)

Q

Quicklime, 1:153

R

Radiation exposure, 3:652-653
 Raschig, Friedrich, 2:561
 Rayon, 1:207-210
 Recreational drugs
 amyl nitrite, 1:89-82
 butane, 1:121
 hexane, 2:355
 nicotine, 2:489-490
 nitrous oxide, 2:514, 2:516
 Red mercury(II) sulfide, 2:439-442
 Refrigerants, 1:180, 277-281
 Regnault, Henri Victor, 1:189
 Reichstein, Tadeusz, 1:95-96
 Reimer-Tiemann reaction, 3:875
 Remsen, Ira, 3:689-690
 Repe, Walter, 1:28
 Retinol, 3:677-681, 678 (ill.)
 Riboflavin, 3:683-687, 684 (ill.)
 Ritthausen, Karl Heinrich, 2:465
 Roche method, 1:111
 Rochow, E.G., 2:593
 Roebuck, John, 3:826
 Röhm, Otto, 2:583
 Röhm and Haas, 2:583, 2:585
 Rotschy, A., 2:487
 Rouelle, Hilaire Marin, 3:867
 Rowland, F. Sherwood, 1:190, 278
 Rubber, synthetic
 isoprene, 2:383-384
 1,3-butadiene, 1:1-4
 Poly(styrene-butadiene-styrene), 2:571-574
 polystyrene used in, 3:598

polyurethane, 3:609-613
Rufen, 1:9-10
Rum, 3:810
Runge, Friedlieb Ferdinand, 2:559
Rust, 2:371

S

Saccharin, 3:689-693, 690 (ill.)
Sachs, Julius von, 1:219
Sal ammoniac, 1:63-64
Sal soda, 3:729-733
Salt and salt substitutes, 2:465-468; 3:641, 653, 735-739
Saltpeter, 3:655-658
Saponification, 2:349-350
SBS rubber, 2:571-574
Schaefer, Vincent, 3:702-703
Scheele, Karl Wilhelm, 1:235; 2:349, 392, 525; 3:629, 760
Schlatter, James M., 2:401-402
Schnell, Hermann, 2:576
Schönbein, Christian Friedrich, 1:201
Scrubbers, 1:153
Scurvy, 1:93-95, 240
Sea water, 3:639-640, 736
sec-butyl acetate, 1:125-127
sec-butyl mercaptan, 1:129-132
Semiconductors, 1:247-250
Semon, Waldo Lonsbury, 3:616-617
Silica gel, 3:696-697
Silicon dioxide, 3:695-700, 696 (ill.)
Silicones, 2:591-596
Silver iodide, 3:701-704, 702 (ill.)
Silver nitrate, 3:705-709, 706 (ill.)
Silver oxide, 3:711-713, 712 (ill.)
Silver(I) sulfide, 3:715-717, 716 (ill.)
Simon, Eduard, 3:597
Simpson, Sir James Young, 1:211
Sinn, Hansjörg, 2:581, 582
Skin treatments
 petroleum jelly, 2:548-549
 phenol, 2:562
 sunscreens, 3:886
 talc, 2:425
 tretinoïn, 3:680
Skunk spray, 1:130-131
Smokeless gunpowder, 1:204
Soaps and detergents, chemicals in
 phosphates, 2:568
 potassium hydroxide, 3:648
 sodium hydroxide, 3:756
 sodium tetraborate, 3:791-792
 triclocarban, 3:861
 triclosan, 3:864
 See also Cleaning products, chemicals in
Soapstone, 2:423-427
Sobrero, Ascanio, 2:507
Soda ash, 3:729-733
Sodium acetate, 3:719-722, 720 (ill.)
Sodium benzoate, 1:107
Sodium bicarbonate, 3:723-727, 724 (ill.)
Sodium carbonate, 3:729-733, 730 (ill.)
Sodium chloride, 3:735-739, 736 (ill.)
Sodium cyclamate, 3:741-745, 742 (ill.)
Sodium fluoride, 3:747-751, 748 (ill.)
Sodium hydroxide, 3:648, 753-757, 754 (ill.)
Sodium hypochlorite, 3:759-763, 760 (ill.)
Sodium perborate, 3:765-768, 766 (ill.)
Sodium percarbonate, 3:767
Sodium perchlorate, 2:541-543
Sodium phosphate, 3:769-772, 770 (ill.)
Sodium polyacrylate, 3:773-777, 774 (ill.)
Sodium silicate, 3:779-783, 780 (ill.)
Sodium sulfite, 3:785-787, 786 (ill.)
Sodium tetraborate, 3:789-792, 790 (ill.)
Sodium thiosulfate, 3:795-798, 796 (ill.)
Soil additives
 calcium hydroxide, 1:144, 148
 calcium sulfate, 1:167
 copper(II) sulfate, 1:252
 gamma-1,2,3,4,5,6-hexachlorocyclohexane, 2:335
 potash, 3:635
 See also Fertilizer, chemicals in
Solvay, Ernest, 1:64; 3:725, 731
Solvents
 butyl acetate, 1:126
 butyl mercaptan, 1:131
 dimethyl ketone, 1:290
 ethyl acetate, 1:293-294
 hexane, 2:353-354
 isopropyl alcohol, 2:388

Soubeiran, Eugene, 1:211
 Specht, Walter, 2:407
 Stahl, Georg Ernst, 1:23
 Stalagmites and stalactites, 1:143, 145
 Standard Oil Co., 2:387
 Stannous fluoride, 3:799-802, 800 (ill.)
 Stannous hexafluorozirconate, 3:801
 Steam cracking, 1:2
 Steel production, 1:152
 Stockings, nylon, 2:522
 Stroke, protection against, 1:33-34
 Styrene, 2:303-304; 3:803-806, 804 (ill.)
 Styrofoam, 3:599
 Sucaryl, 3:743
 Sucrose, 2:329, 2:331; 3:807-811, 808 (ill.)
 Sucrose polyester, 3:813-817, 814 (ill.)
 Sugars
 fructose, 2:329-332
 glucose, 2:343-347
 high-fructose corn syrup, 2:330-331
 sucrose, 2:329, 331; 3:807-811
 Sulfur dioxide, 3:819-823, 820 (ill.)
 Sulfuric acid, 3:825-829, 826 (ill.)
 Sunscreens, 3:886
 Super Glue, 1:259-263, 261-262
 Supercritical carbon dioxide, 1:139, 179
 Sveda, Michael, 3:741
 Swallow, J.C., 2:580
 Sweeteners, artificial
 glycerol, 2:349-352
 L-aspartyl-L-phenylalanine methyl ester, 2:401-405
 saccharin, 3:689-693
 sodium cyclamate, 3:741-744
 sucrose, 3:807-811
 Szent-Györgyi, Albert, 1:95; 3:673

T

Tachenius, Otto, 3:659
 Takaki, Kanehiro, 3:849
 Talbot, William Henry Fox, 3:706
 Talc, 2:423-427
 Tannic acid, 3:831-835, 832 (ill.)
 Tanning (hides), 3:831-834
 Tarnish, 3:715, 717

Tastes, bitter, 1:271-274
 Tea, 1:137-141
 Teflon, 3:603-607
 Terpenes, 2:381-385
tert-butyl acetate, 1:125-127
tert-butyl mercaptan, 1:129-132
 Testosterone, 3:837-841, 838 (ill.)
 T&H Smith, 1:271
 Thénard, Louis Jacques, 1:116; 2:363
 Theobromine, 1:138-139; 3:843-846, 844 (ill.)
 Thiamine, 3:847-851, 848 (ill.)
 Thiele, Johannes, 1:101
 Thyroid gland, 3:652-653
 Tilden, Sir William Augustus, 2:384
 TNT, 1:15-18; 3:855-856
 Tobacco, 2:487-491
 Tocopherols, 1:37-40
 Tofu, 2:413
 Tollens, B.C.G., 3:855
 Toluene, 1:105-107; 3:853-857, 854 (ill.)
 Tooth decay, 3:748-749, 799-802, 811
 Torches, 1:27-29
 Torpex, 1:17
 Tournefort, Joseph, 1:64
 Tribasic calcium phosphate, 1:155-159
 Tribasic sodium phosphate, 3:769-772
 Triclocarban, 3:859-862, 860 (ill.)
 Triclosan, 3:863-866, 864 (ill.)
 Troostwyk, A. Paets van, 2:307
 Tschunkur, Eduard, 2:572
 TSP, 3:771
 Tsvett, Mikhail, 1:219
 Twenty-Mule-Team Borax, 3:791-792
 2-(4-Isobutylphenyl)-propionic acid, 1:9-13, 10 (ill.)
 2,4,6-Trinitrotoluene, 1:15-18, 16 (ill.)
 2,2-Dichlorodiethyl sulfide, 1:5-8, 6 (ill.)
 Tylenol, 1:21

U

Umetaro, Suzuki, 3:847
 Upjohn Co., 1:10-11
 Urea, 3:867-871, 868 (ill.)

V

Vaccines, 1:47

Vanillin, 3:873-877, 874 (ill.)

Vaseline, 2:547-549

Vasodilators, 1:89-90

Vermillion, 2:440

Vicks VapoRub, 1:173; 2:437-438

Vinegar, 1:23-26

Viscose Corp. of America, 1:208

Vision, 3:679

Vitamins

alpha-tocopherol (E), 1:37-40

ascorbic acid (C), 1:93-97

beta-carotene, 1:109-113; 3:678, 3:680

cyanocobalamin (B₁₂), 1:265-268

folic acid (B₉), 2:321-324

niacin (B₃), 2:483-486

pyridoxine (B₆), 3:673-676

retinol (A), 3:677-681

riboflavin (B₂), 3:683-687

thiamine (B₁), 3:822, 847-851

See also Minerals (nutritional)

Volpenhein, Robert, 3:813

Vonnegut, Bernard, 3:703

W

Wackenroder, Heinrich Wilhelm

Ferdinand, 1:110

Waldie, David, 1:211

Warning odors, 1:87

Washing soda, 3:729-733

Water, 3:879-884, 880 (ill.)

Water treatment

 aluminum potassium sulfate, 1:55

ammonium sulfate, 1:78

fluoride, 1:43; 3:748-749

sodium sulfite in, 3:786

softeners, 3:737

Wedgwood, Thomas, 3:706

Wehmer, Carl, 1:235

Wells, Horace, 2:514

Whipple, George Hoyt, 1:266

Whitehall Laboratories, 1:11

Wiegleb, Johann Christian, 2:525

Wieland, Heinrich Otto, 1:224-225

Wilbrand, Joseph, 1:15

Williams, Robert R., 3:849-850

Willow bark, 1:31-32

Wills, Lucy, 2:321

Winhaus, Adolf, 1:224-225

Wislicenus, Johannes, 2:392

Wöhler, Friedrich, 1:27; 2:525; 3:867

Wood alcohol, 2:449-453

Wood preservatives, 1:248, 252

Woodward, Robert Burns, 1:219, 225, 267

Wounds, sealing of, 1:260-261

Wulff process, 1:28

Wurtz, Charles Adolphe, 2:313, 317

X

Xylene, 3:855

Z

Zeidler, Othmar, 1:283

Zeolites, 1:257

Ziegler, Karl, 2:384, 580, 2:582, 588-589

Zinc oxide, 3:885-888, 886 (ill.)