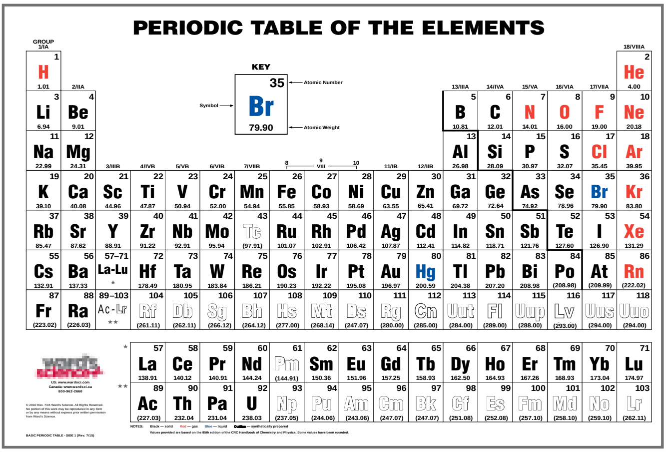
All About Elements:

Fluorine

Ward's All About Elements Series

Building Real-World Connections to the Building Blocks of Chemistry



The periodic table of elements is an essential part of any chemistry classroom or science lab, but have you ever stopped to wonder about all of the amazing ways each element is used to create the world around us? Each of the trillions of substances in our universe can be tied back to just these 118 simple, yet powerful elements.

In our ***All About Elements*** series, we've brought together the most fascinating facts and figures about your favorite elements so students can explore their properties and uses in the real world and you can create chemistry connections in your classroom and beyond.

Look for a new featured element each month, plus limited-time savings on select hands-on materials to incorporate these element in your lessons.

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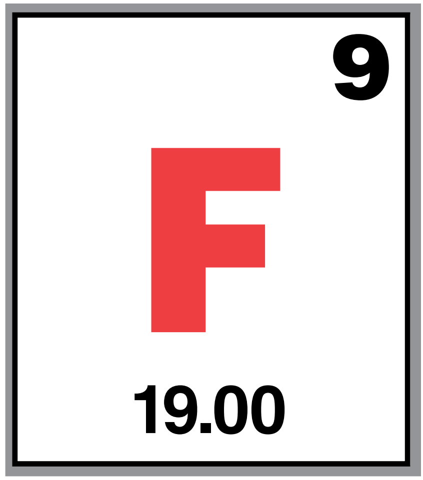
Fun Facts About... Fluorine



1. Fluorine has both toxic and helpful properties that make it widely used, from medical treatments to poisonous gases and nuclear applications.

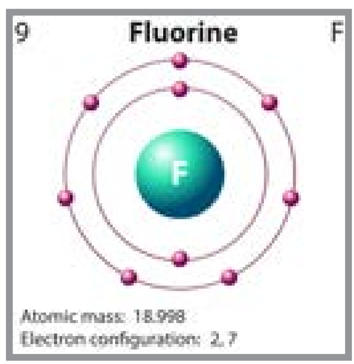
2. Fluorine gets its name from the Latin word "fleure" which means "to flow", which is what fluorine allows metals to do. [1]

3. Although the element fluorine was known to exist in 1813, it wasn't until 1886 that Henri Moissan was able to isolate the element after numerous failed, and sickening, attempts.

4. Fluorine wasn't mass produced until it was needed for nuclear applications during World War II.

5. One of the reasons you've probably had only a cavity or two is due to the fluoridation of water, which was first tested in 1944 in Grand Rapids Michigan. Here it was determined that fluoride helped strengthen teeth.

All About Fluorine:



Fluorine is the ninth element on the periodic table found in group 17 (7A) with an atomic number of 9 and a symbol F. Fluorine gets its name from the Latin word "fleure" which means "to flow", which is what fluorine allowed metals to do. [1] Fluorine came also from the "fluor" in fluorspar (which we now know to be calcium fluoride.)[1] Group 17 is known as the halogens and fluorine is known to be the most reactive of all of the halogens, gaining one electron in most cases to become a stable fluoride ion. Fluorine is the smallest of the halogens with an electron configuration of $1s^2 2s^2 2p^5$, and has both the highest ionization energy in its group and the highest electronegativity on the periodic table. Fluorine has a density of 1.696 grams per liter which is the lowest density of all of the elements in its group and is the second least dense gas, behind neon with a density of 0.9 g/L, in its period on the periodic table. [4]

Before fluorine was isolated in 1886, it was known to exist in a number of compounds.[1] These compounds are where you will find fluorine most of the time today, as diatomic elemental fluorine doesn't exist in nature, but rather must be created by electrolysis. Some of the more common compounds of fluorine are fluorspar, hydrofluoric acid, potassium fluoride, sodium fluoride, tin (II) fluoride, sodium monofluorophosphate, uranium hexafluoride, dichlorodifluoromethane, and polytetrafluoroethylene.[5] These compounds range in use from additives to toothpaste for strengthening teeth, to use in atomic bombs for their explosive nature. Fluorine has also been found to react with the noble gases xenon, krypton and radon despite their full octet of electrons. [6] Although it is found in these and other compounds, fluorine typically only has an oxidation state of -1 due to its high electronegativity. In addition, there is only one stable isotope of fluorine, F-19 with 9 protons and 10 neutrons.[5]

Properties of Fluorine

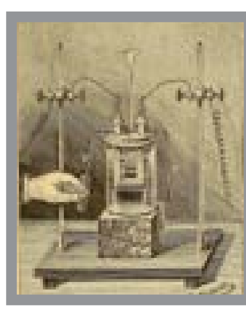
In its pure gaseous state, fluorine is pale yellow and extremely corrosive reacting with most organic and inorganic substances.[6] Specifically, fluorine's color can only be seen by looking down at it through a test tube, not from the side.[18] In addition, it can be lethal and will react with glass, ceramics, carbon, finely divided metals and even water to create an explosion.[6] Because of its explosive nature, it is commonly shipped as a cryogenic liquid.[10] If fluorine is inhaled or absorbed through the skin it is toxic, but even at lower than lethal concentrations it can cause chemical burns.[10] Some of these properties have been put to good use, as hydrofluoric acid is used to etch glass, something that is a delicate process.

Even though fluorine is found in trace amounts on Earth (it is the 13th most common element in Earth's crust,[11] and only amounts to 1.3 milligrams per liter in the ocean[5]) and in the universe, it is most commonly found in the form of a fluoride ion. For instance, the fluoride ion is present at about 3 ppm in the human body![7] Compounds that are added to toothpaste and drinking water are all in the form of the ion, not the toxic gas. All of the stable naturally occurring fluorine is F-19, with isotopes used for a variety of medical purposes and having a variety of half lives. The half lives range from 4.557 x 10^{-22} seconds for F-15 to 109.77 minutes for F-18 making the radioisotopes quickly removed from the body.[9] This isotope with 9 protons and neutrons, F-18, is commonly used as a radioisotope to detect cancer, for cardiac and brain imaging as well as monitoring the progress of treatment of these diseases because it has been "proven to be the most accurate non-invasive method of detecting and evaluating most cancers."[8].

Discovery and History

If you thought, based on some of the properties described, that fluorine had a rather interesting history you'd be correct! First, Fluorine is created in stars and quickly breaks down producing oxygen and helium or neon and hydrogen after undergoing fusion with hydrogen or helium respectively [16]. Although it wasn't known exactly how it worked, the first use of the mineral fluorite (made from fluorine and calcium) was for flux [12]. Flux is introduced when smelting ores and allows the metal to flow while it removes the impurities, or slag [12]. This is part of where the name came from for fluorine, as fleure means to flow which is what this mineral allowed metals to do. The first record of fluorine is from 1670 in a set of instructions on how to etch glass with Bohemian emerald, now known to be CaF_2 [5]. Specifically the chemist Schwandhard found that glass could be etched when it was exposed to fluorspar treated with acid.

Early 1800's



Moissan's fluorine cell, from his 1887 publication

In the early 1800s numerous chemists carried out experiments on fluorspar (now known to be calcium fluoride) including Gay Lussac, Humphry Davy and Joseph Priestly. In fact, none of them were able to isolate fluorine, but many produced hydrofluoric acid, which unfortunately led to fatalities and to blindness.[1] Davy then used the hydrofluoric acid and became ill trying to isolate fluorine. Davy and a French scientist, Andre-Marie Ampere, exchanged letters on how the acid may have a new element in it and the element, although not yet isolated, was given its name of fluorine by Ampere.[1] Finally in 1886, a French chemist named Henri Moissan was able to isolate fluorine using electrolysis of dry potassium hydrogen fluoride in anhydrous hydrofluoric acid contained in a platinum container that he cooled to -31oC.[1] Moissan's work was interrupted four times due to poisoning caused by the fluorine he was trying to isolate, but his work was awarded in 1906 with the Nobel Prize in chemistry. [1] It wasn't until World War II that commercial production of fluorine began to help with nuclear energy and the bomb project. [6]

Fluoridated Water

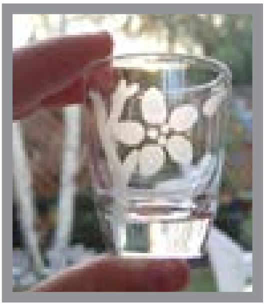
In addition to elemental fluorine, fluoridation of water and compounds containing fluorine have interesting histories. Fluoridated water came to be after researchers in 1901 in Colorado Springs noticed that people had black and brown spots all over their teeth in addition to "mottled tooth enamel."[13] They traveled the country to other sites where people lived with cracked teeth trying to determine what might be causing the problem. In 1923 the problem was found to be the water source, but it wasn't until 1931 when photospectrographic analysis was performed on samples of water that the problem was realized.[13] It was also known that the mottled tooth enamel was resistant to decay. Dr. Trendley Dean wondered if fluoride added to drinking water at lower levels could help fight decay. After a few years of research and in discussions with the city commission of Grand Rapids Michigan, fluoride was added to the water there in 1944. The tooth decay rate of over 30,000 children that were studied decreased by 60 percent.[13] According to the CDC, most additives that are used in the US come from phosphorite rock that is treated with sulfuric acid to produce a phosphoric acid- gypsum slurry.[14] Once heated, this slurry releases both hydrogen fluoride and silicon tetrafluoride gases which are captured and then condensed into water to use as fluoride ions.[14]

Where in the world is Fluorine?

Fluorine is only found in Earth's crust in an abundance of 0.054%[2] and in the universe as 0.00004% abundance.[3] In Earth's crust it can be found in coal, clay and rocks, as well as in the soil with approximately 330 parts per million on average. [16] Fluorite, which is the most common form of fluorine, is found in numerous locations around the world including Argentina, Canada, China, France, Mexico, Russia, Spain and the United States.[15] That being said, the primary areas for mining are China, Mexico and western Europe.[16] Depending on where the mineral containing fluoride is collected, it will fluoresce different colors. Fluorite is one of the more popular minerals in the world among mineral and gem collections, despite its toxicity when treated and turned into hydrofluoric acid or fluorine gas.[15] Fluoride ions are added to water in the United States so that the concentration is about 0.7 milligrams per liter and because fluoride can be natural in some places (coming from minerals) the maximum level allowed in 4.0 milligrams per liter.[14]

Compounds containing fluorine also exist in the atmosphere, depleting the ozone as CFC's or chlorofluorocarbons. These were originally used as propellants in many spray cans are harmless to humans. Sadly, it was later discovered that the chlorine reacts with ozone, O_3, removing the protective layer from the earth's atmosphere. Since 1996, CFC's have been banned and hydrofluorocarbons (HFCs) have been created and used as replacements. These compounds are specifically used in refrigerators, air conditioning systems and as fire retardants. HFCs are believed to be safer than CFCs for the ozone layer and are more controlled in terms of where they are used.[17]

Uses of Fluorine Today:



Elemental fluorine is so highly reactive and toxic to humans that it is rarely used today. It can even react with diamond, one of the strongest materials on Earth. Compounds containing fluorine are much more common and useful than diatomic fluorine gas in its normal state. For example, gem collectors search the globe to obtain different colors of nephrite, topaz and feldspar each of which contains fluorine components. Hydrofluoric acid, although it is a weak acid, is highly toxic to humans, but is used in etching and frosting glass. Fluorite can be used in the processing of metals and smelting of iron. Another very important use of fluorine is in refrigerants for both air conditioners and refrigerators that have be changed since CFS's were banned to other fluorine containing compounds. Below you will find other applications of fluorine.

Teflon:

Teflon, although commonly thought of today as toxic, was once used on all non-stick pans. It is produced, along with other fluoropolymers, around 180,000 metric tons annually.[16] Teflon, Polytetrafluoroethene, is considered a high temperature plastic that doesn't melt until 326°C (or 620.3°F for those of us used to baking at temperatures upward of 400°F). In addition to being used in cooking, it is also used in cable insulation, plumber's tape and as the base of Gore-Tech® which is used to waterproof shoes and clothing.[7]

http://home.howstuffworks.com/nonstick-cookware.htm

Nuclear Power and Bombs:

Uranium hexafluoride, UF6, is helpful in both nuclear reactors and bombs in processing nuclear fuel. UF6 can be used in all three states of matter: solid for storage, liquid for filling or emptying containers or equipment and as a gas for processing.[19] The uranium that is used has to be enriched so the percent of uranium 235 is higher. The UF6 helps this process take place by enriching one sample and depleting the other sample effectively "transferring" the U-235 from one sample to another.[20]

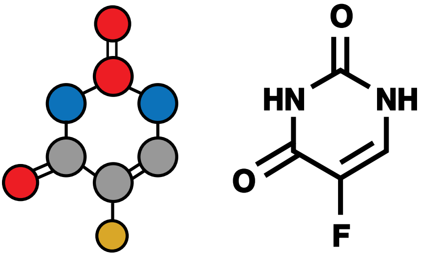
Breaking Bad: Dissolving Bodies?

If you've ever seen the popular T.V. series, *Breaking Bad*, you know Walter White and his partner in crime Jesse Pinkman. In one episode, they tried to dispose of a rival's body in a bathtub using hydrofluoric acid and ran into some trouble. Instead of pouring the contents into a plastic container (which hydrofluoric acid CAN'T dissolve), Jesse pours the HF into a bathtub, which is then dissolved by the acid. Later, the tub and disintegrating body fall through the ceiling onto the floor below. While there is some truth to this concept, in reality, it is unlikely that the tub would go through the ceiling in this manner. The episode was even debunked by Mythbusters, which found that after eight hours of soaking in HF, all the materials that would have been between the person they were trying to dispose of and the next floor, had deteriorated to different degrees, but none were completely destroyed[21].

http://www.bbc.com/news/magazine-23710654

Chemotherapy:

The drug Fluorouracil is used in chemotherapy treatments for cancers ranging from breast and colon to cervical and head cancer. It can also be used as a topical cream on basal cell cancer.[22] This specific chemotherapy drug is considered an antimetabolite, which is very similar to substances normally in cells and therefore easily integrated. Fluorouracil interferes specifically with DNA and RNA synthesis by "mimicking the building blocks necessary for synthesis."[22]



Other Pharmaceuticals:

In addition to being used in chemotherapy, which is known for its adverse side effects, fluorine containing compounds are used in numerous other medicines. The element appears in approximately 20% of all pharmaceuticals mainly due to the strength of the bonds it forms with other atoms—specifically carbon.[7] Since fluorine is so reactive, it isn't found alone in nature and the bonds that it forms are strong due to fluorine's high electronegativity. Many times fluorine is used to protect the drug from degradation helping to last longer and extending the active lifetime inside the patient's' body.[7] To do this, scientists replace hydrogen with fluorine creating a stronger bond between elements. In addition, fluorine can help change the shape of the molecule it is introduced in helping it bind better to the protein it is targeting.[7] Common drugs that use these properties of fluorine include Prozac, an anti-depressant, and Lipitor, a cholesterol-lowering drug.[7].

Rocket Fuel?

Elemental fluorine, F2, has been studied as a possible rocket propellant.[6] Engines that run on liquid fluorine have been developed and fired successfully, but are not widely used due to safety concerns.[23] For example, fluorine is extremely toxic; It is considered a "super oxidizer" and it reacts violently with almost anything except substances that have already been fluorinated, lighter noble gases and nitrogen.[23] Liquid fluorine has also been mixed with liquid oxygen to improve the performance of LOX- burning engines to create FLOX (Fluorine-Liquid-Oxygen burning). Despite its ability to propel and explode, due to safety concerns around toxicity, fluorine has been abandoned by most nations hoping to travel to space.

Toothpaste:

Pearly whites. Chompers. What have we to thank for these novel nicknames other than toothpaste. Fluoride-containing compounds are added to toothpaste to help strengthen the enamel and prevent cavities similar to how fluoride is added to water.[24] When toothpaste is applied at the tooth buds, the hydroxyapatite enamel of the tooth is replaced by fluorapatite. This compound, fluorapatite, is more resistant to decay than hydroxyapatite because it is less soluble in acid than hydroxyapatite[24].

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Ward's® Jade (Nephrite) Hand Specimen

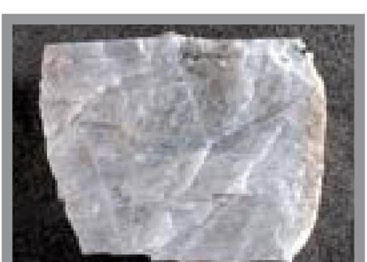
Amphibole group; green, sawed sections, compact variety of actinolite..

Item Number: 460979

Ward's® Topaz (Crystals)

Gemmy, sherry-colored crystals and sections approximately ⅜–½".

Item Number: 468220



Ward's® Oligoclase

Feldspar group; white-to-gray cleavages of sodic plagioclase; some show twinning. Specimens range in size from 2" x 3" to 3" x 4".

Item Number: 465804

Sodium Fluoride

NaF

F.W.: 41.99 CAS#: 7681-49-4

Hazard: Toxic

Shelf Life (months): 36

Storage: Blue

Soluble: Water and Alcohol

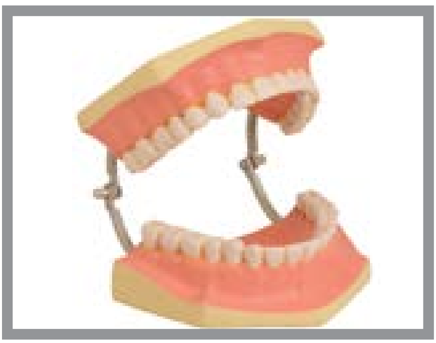
bp (°C): 1700

mp (°C): 993

Density (g/mL): 2.78

Grade: Reagent, Powder

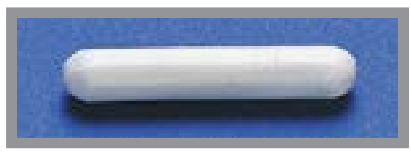
Item Number: 470302-540



Eisco® Dental Care Model

Giant set of teeth and gums mounted on hinges, excellent tool for showing proper dental care.

Item Number: 470136-058



Teflon® Stirring Bar

Teflon coating makes each bar chemically inert for high purity contact. $1\frac{5}{8}$"x $\frac{5}{16}$".

Item Number: 181707

Calcium Fluoride

CaF_2

F.W.: 78.08 CAS#: 7789-75-5

Hazard: Slightly Toxic, Irritant

General Storage

Shelf Life (months): 36

Storage: Green

Soluble: Ammonium Salts

bp (°C): Decomposes

mp (°C): 1423

Density (g/mL): 3.180

Grade: Laboratory, Powder

Item Number: 470302-540

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