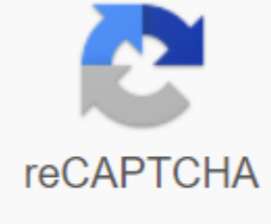




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Ionic charge of fluoride

Octet rules are usually used to describe the charges that group I, II, VI, and VII atoms have in ion compounds by saying that noble gases are stable and have octavalent electrons. This is true, but it is important to understand why the noble gas configuration is stable, and it seems magical that some atoms do not adopt noble gas configurations in ionic compounds without knowing it. Octet rules are a convenient way to predict ion rates quickly and write Lewis structures, but they are not the magic laws that atoms follow. They don't want or try to get a particular electronic structure, but we often talk like them. What really matters is the total energy of electronic composition. The electronic configuration with the lowest energy is the most stable electron. The total energy of the electron composition, where stability really means lowest energy, is governed by an effective nuclear charge in which the energy of the electron orbit is seen by electrons. That is, orbital energy determines the order in which atomic orbits are transferred by electrons in the ground state. We know that from experienced evidence and quantum machine calculations, order is carried out in this way for most elements (rather than some transition metals and heavy elements): electrons occupy these trajectories from the ground (called the Aufbau principle). Therefore, in the fourth period element, it is known that the $3d$ subshell is filled before the $4p$ subshell does it. An atomic orbit diagram indirectly tells us something else. The main quantum number is n dollars (equal to the number of rows or periods in the case of valence shells), which corresponds greatly to the distance from the nucleus of the highest electron density region of the orbit. In other words, a subshell of n to 4 dollars will have most of the electron density away from the nucleus than the subshell of n This means that the electrons are closer to the nucleus than the ones in the first generation, even though the electronics are the first to have a high energy. Electrons located almost proportional to the distance of electrons from the nucleus have a negative charge, and since the nucleus is positively charged, there is an electrostatic potential energy between the electrons and the nucleus. If electrons were uniformly distributed throughout the atom, the nuclear charge that each electron felt is expected to be equal to the number of protons (the number of atoms). However, due to the rules of quantum mechanics (the solution allowed for Schrodinger's equation), electrons are not uniformly distributed and occupy orbit in certain arrangements. Some electrons are closer n to the nucleus than others, so this is partially screened by internal electrons and the outer electrons experience somewhat less effective nuclear charge than the full nuclear charge. This diagram of Wikipedia shows the idea using the Bohr model (this is not an exact image of what the electrons in orbit are lying, but only the relative distance from the nucleus is correct), and when valence electrons experience a large effective nuclear charge, it is difficult to remove (increases the energy cost to remove them) and the effective nuclear charge is low, making it easier to remove. It is very difficult to accurately calculate the actual effective nuclear charge, but it is very easy to estimate - all you need to do is take the number of protons and then remove the number of screened electrons. Basically, we're assuming that each electron between a valence electronic shell and a nucleus cancels one proton $Z - Z_{\text{inner}}$. Z the number of dollars, all you have to do is count the number of electrons that are less than the valence shell . For example, if you take fluorine, the electronic configuration is as follows: Z Compare it to lithium: a valence electron of fluorine is about seven times harder to remove than a lithium valence electron. In fact, any element on the left side of fluorine will give up electrons easier than fluorine, with a low effective nuclear charge of the valence electron shell. In other words, rather than fluorine not giving up electrons, nothing can take them out of it. What happens if you add an electron? Z This means that fluorine is energetically advantageous to pick up electrons because additional electrons still see large positive nuclear charges. What happens if I add another one to make it Z ? Note that the number of screening electrons has jumped to 10 because the new valence shell has a value of n In other words, due to the allowed orbital configuration, new electrons would see negative charges on the nucleus - it needs energy to maintain it. If you process this process with all of the fixed charge elements, you can find similar things - an effective nuclear charge will be negative at that time until you get too much (octets), large for the elements that get electrons. In the case of fixed-charge metals, something else happens. Let's in magnesium. The remaining value jumps to +10 if you remove the two electrons. The same is true for all Group I and II metals - once the Noble gas core (octet) is reached, the effective nuclear charge seen by valence electrons jumps. This explains why the fixed charge element always has the same charge in ionic compounds, but what about transition metals? The electronic configuration looks like this: the one with the highest value of the electronics is the one with the highest value of the electronics, which ones are the highest Note that the effective nuclear charge they see is If you get rid of the electrons, you'll get rid of them. $Z - a_0 - a_1 - a_2 - \dots - a_n$ - has not changed. You can delete and get another one to create the Now that the definition of valence is a little clearer, and we know that the maximum value corresponds to the valence shell, n is the orbital shape? d orbit has been found to be greater than the orbit of s and p , so Electrons of s and p will screen the nucleus from d electrons. This means that d -electrons are then removed, and the effective nuclear charge seen by them is This is a larger charge, but other elements (especially oxygen and halogen) are not large enough to separate d electrons yet. This gives a range of chromium charges (oxidation states) from +1 to +6. For Cr , the electronic configuration will look like this: As you can see, the rules of quantum mechanics determine the relative size of the atomic orbits and the order in which they are expected, and then determine which effective nuclear charges valence electrons see for a particular medium atom or ion. The combination of the two results results in charge constraints that atomic ions can have, the actual results, group I, II, VI, and VII elements adopt a filled octet or noble gas configuration, other groups have more flexibility. This article describes fluoride ions. For the study of fluorine compounds, see Fluorine Compounds. See Fluoride for fluoride additives used in toothpaste. Don't confuse it with fluorid or fluorine. Fluoride Name IUPAC Name Fluoride[1] Identifier CAS No. 16984-48-8 Y 3D Model (JSmol) Interactive Image ChEBI CHEBI:17051 ChemBL CHEMBL1 362 Y Chem Spider 26214 Y Gumelin Reference 14905 KEGG C00742 Y MeSH Fluoride Pubchem CID 28179 UNII Q80VPU408O Y Comptox Dashboard (EPA) DTXSID9049617 InChI -M Y SMILES [F-] characteristic chemical formula F-molar mass 18.998403163 g-mol-1 hydrogen fluoride hydrogen conjugated chemical Std MOLARENTPy (So298) 145.58 J/mol K (Tm)[2] Formation of Std Entharpy ((Delta) H \ominus 298) - 333 kJ mol-1 Related Compounds Other anionic chloride chlorochroide iodide unless otherally described, data on the material of the standard state is given (25 . FactBox refers to fluoride (ˈfluəridə , ˈflɔːr-ɪ)^[3] is an inorganic, Monia anion with a chemical formula F⁻ (also written [F⁻]) in which salt is typically white or colorless. Fluoride salts typically have a unique bitterness and are odorless. Its salts and minerals are important chemical reagents and industrial chemicals mainly used in the production of hydrogen fluoride. Fluoride is classified as a weak base because it is only partially associated with the solution, concentrated fluoride is corrosive and can attack the skin. Fluoride is the simplest fluorine anion. In terms of charge and size, fluoride ions are similar to hydroxide ions. Fluoride ions occur on Earth in some minerals, especially fluorine, but only in trace amounts in natural waters. The nomenclature, and compounds containing ionic fluoride, include compounds fluoride does not solve. Nomenclature does not distinguish between these situations. For example, sulfur hexafluoride and carbon t fluoride are not sources of fluoride ions under normal conditions. The strain name fluoride is a valid IUPAC name, it is determined according to the additive nomenclature. However, the name fluoride is also used in the composition IUPAC nomenclature that does not take into account the nature of the binding. Fluoride is also used non-systematically to describe compounds that release fluoride during dissolution. Hydrogen fluoride is itself an example of a non-systematic name of this property. However, it is also a trivial name, and the IUPAC name of fluoran is also preferred. [Need quote] generated fluorine crystal fluorine is estimated to be the 13th richest element in the Earth's crust and is widely dispersed in nature in the form of fluoride. Many minerals are known, but the most commercially important is fluorine (CaF2), which is about 49% fluoride. Soft and colorful minerals are found all over the world [water fluoride is naturally present at the freshest and lowest concentrationsEspecially in urban areas, as well as the source of rainwater. [5] Seawater levels of fluoride typically range from 0.86 to 1.4 mg/L, with an average of 1.1 mg/L[6] (milligrams per liter). For comparison, chloride concentration in seawater is about 19 g / L. Low concentrations of fluoride, insoluble alkaline fluoride, for example, reflects CaF2. The concentration in fresh water is much different. Surface water, such as rivers and lakes, is typically contained between 0.01 and 0.3 ppm. [7] Groundwater (well water) concentration varies further depending on the presence of local fluoride-containing minerals. For example, natural levels below 0.05 mg/L were detected in parts of Canada, but up to 8 mg/L in some parts of China. In some areas, such as Tanzania and parts of India, it generally rarely exceeds 10mg/liter[9], drinking water contains dangerously high levels of fluoride, causing serious health problems. Globally, 50 million people naturally receive water from water supplies close to the optimal level. In other places the level of fluoride is very low, sometimes causing fluorination of the public water supply, bringing the level to about 0.7 to 1.2 ppm. Fluoride is present in the rain and can significantly increase concentrations when exposed to volcanic activity and air pollution caused by the burning of fossil fuels and other types of industries. [11] In plants all vegetation contains fluoride absorbed from soil and water. [8] Depending on the plant, concentrating fluoride from the environment than other plants. All tea leaves contain fluoride. However, mature leaves contain 10-20 times the fluoride levels of young leaves from the same plant. [13] [14] [15] Chemical properties Basic fluoride can act as basic. Proton (H +): can be combined with F⁻ + H + HF \rightarrow (1) (1) the mesothermal reaction to form hydrogen fluoride (HF) is a co-acting acid fluoride. In aqueous solution, fluoride has a pKb value of 10.8. Therefore, the base is weak, there is a tendency to remain as fluoride ions rather than generate a large amount of hydrogen fluoride. That is, equilibrium supporting the left side of the water, F⁻ + H2O \rightleftharpoons . Fluoride is clear in this regard among halides. The sameness of the solvent has a dramatic effect on shifting the equilibrium to the right side, greatly increasing the decomposition rate. Structure of fluoride salts Many salts containing fluoride adopt a myriad of structures. Typically, fluoride anions are surrounded by four or six cations, as is typical of other halides. Sodium fluoride and sodium fluorideAdopt the same structure. For compounds containing a pluralation of fluoride per cation, as indicated by the main fluoride fluorine (CaF2) Ca 2 + ions are surrounded by eight F⁻ cores, the structure often deviates from the structure of chloride. In CaCl2, each Ca2+ ion is surrounded by six Cl-centers. Difluoride of transition metals often adopts a rutile structure, dichloride has a cadmium chloride structure. Inorganic chemistry When processed with standard acids, fluoride salts are converted into hydrogen fluoride and metal salts. With strong acid, you can double-proton to give H2F+. Oxidation of fluoride gives fluorine. The solution of inorganic fluoride in water contains F⁻ and bifluoride HF-2. [16] Inorganic fluoride includes a small number of fluoride easily soluble in water without receiving significant hydrolysis. In terms of its reticability, fluoride is significantly different from chlorides and other halides, it is more strongly solvated in a protic solvent for a smaller radius / charge ratio. Since both have the same geometric shape, its closest chemical relativity is hydroxide. The most fluoride salts of naked fluoride dissolve to give bifluoride (HF2-) anions. The source of true F-anions is rare because very basic fluoride anions abstract protons from many and even adventurous sources. Non-solvent fluoride present in non-solvents is called naked. Naked fluoride is a powerful Lewis base, [17] and a powerful nucleophile. The fourth-class ammonium salts of naked fluoride include tetramethyl ammonium fluoride and tetrabutylammonium fluoride. [18] Cobalt fluoride senium is another example. [19] However, they are all lacking in structural characterizing in an aprotic solvent. Because of its high basicity, many so-called naked fluoride sources are actually fufluoride salts. In late 2016 imidazolium fluoride was synthesized thermodynamically stable and structurally characterized, which is the closest approximation of the naked fluoride source (acetonitrile) in a non-protic solvent. [20] The demanding

imidazolium cation to testerric stabilizes discrete anions and protects them from hydration. [21] In biochemical physiological pHs, hydrogen fluoride is usually completely ionized into fluoride. In biochemistry, fluoride and hydrogen fluoride are equivalent. Fluorine, in the form of fluoride, is considered to be a micronutrient for human health and is necessary to prevent the tooth cavity and promote healthy bone growth. Tea Plant (*Camelia Sinensis* L.) is a known accumulator of fluorine compounds, it is released when forming an injection such as a general beverage. Fluorine compounds are decomposed into products containing fluoride ions. Fluoride is the most bio-available fluorine, and therefore tea is a potential means for the use of fluoride. [24] Approximately 50% of absorbed fluoridePeriod. The rest can be held in the oral cavity, and in the lower digestive tract. Fasting dramatically increases the absorption rate of fluoride to nearly 100%, and increases dramatically from 60% to 80% when ingested with food. [24] A 2013 study found that the consumption of tea per liter per day could provide a recommended daily intake of 4 mg per day. Some low quality brands can supply up to 120% of this amount. Fasting can increase this to 150%. The study shows that when water fluorination is enabled, the tea-drinking community is at increased risk of dental and skeletal fluorescence. Low-dose fluoride ions in the mouth reduce tooth decay. [25] For this reason, it is used for toothpaste and water fluorination. With much higher doses and frequent exposure, fluoride can cause health complications and be toxic. Applications Related: Fluorine chemical industry, biological aspects of fluorine, fluorine salt and hydrofluoric acid are the main fluorides of industrial value. Compounds with C-F bonds are classified into the region of organofluorine chemistry. The main applications of fluoride, in terms of volume, are in the production of cryolite, Na3AlF6. It is used for aluminum smelting. It used to be mined, but now it is derived from hydrogen fluoride. Fluorine is used on a large scale to separate slag in steel production. The mined fluorine (CaF2) is a commercial product used in steel. Hydrofluoric acid and its anhydrous hydrogen fluoride are also used in the production of fluorine carbon. Hydrofluoric acid has a variety of special applications, including the ability to dissolve glass. [4] Cavity prevention Main article: Fluoride therapy and water fluoride are sold in cavity prevention tablets. Fluoride-containing compounds, for example, used in topical and systemic fluoride therapy to prevent tooth decay, such as sodium fluoride and sodium monofluorophosphate. They are used in many products related to water fluorination and oral hygiene. [26] Originally, sodium fluoride was used fluorine water for fluoride. Hexafluorotic acid (H2SiF6) and sodium hexafluoride is a salt thereof (Na 2SiF6) is a more commonly used additive, especially in the United States. Fluorination of water is known to prevent tooth decay, and the U.S. Centers for Disease Control and Prevention is considered one of the 10 largest public health outcomes of the 20th century. [29] In countries where large concentrated water systems are rare, fluoride is delivered to the masses by fluorinide table salts. Fluorination of water has its critics (see Water Fluorination Controversy). Fluorine toothpaste is commonly used, but is only effective at concentrations greater than 1,000 ppm. [32] Biochemical reagent fluoride salts are biologically commonly usedTreatment that inhibits the activity of phosphates such as serine/sleoneine phosphanase. [33] Fluoride mimics nucleic hydroxide ions at the active site of these enzymes. [34] Beryllium fluoride and aluminum fluoride are also used as phosphate inhibitors, these compounds are structural imitations of phosphate group, it is possible to act as analogs of the transition state of the reaction. [35] [36] Fluoride ion batteries A large team of researchers, including Simon C. Jones of the California Institute of Technology and Christopher J. Brooks of Honda Labs, came up with a liquid electrolyte that reciprocating fluoride ions and was charged at room temperature. [37] The American Institute of Medicine (IAM), which demonstrated its use in [37], updated its estimated average requirements (EA) and recommended diet allowance (RA) for some minerals in 1997 [if there is not enough information to establish EA and RA, an estimate of appropriate intake (AI) was used instead. AI usually matches the actual average consumption, appears to have a need, and is based on the assumption that what people consume will meet that need. Current AI for women over 19 years old is 3.0mg/day (including pregnancy and lactation). AI for men is 4.0 mg / day. AI for children between 1 and 18 years old increases from 0.7 to 3.0 mg/day. The main known risk of fluoride deficiency seems to be an increased risk of the tooth cavity caused by bacteria. Regarding safety, IAM will set an acceptable upper limit intake (ULs) of vitamins and minerals if the evidence is sufficient. In the case of fluoride, UL is 10 mg / day. Collectively, EA, RDR, AI, and ULs are referred to as dietary intake (DRI). The European Food Safety Authority (EFSA) refers to a set of information as dietary standards, showing population reference intake (PRI) rather than RDA and average requirements instead of EAR. AI and UL have the same definition as the United States. For women over 18 years of age, ai is set to 2.9 mg/day (including pregnancy and lactation). The male value is 3.4 mg/day. For children between 1 and 17 years of age, AI increases from 0.6 to 3.2 mg/day with age. These AAs are comparable to 1st class of AI in the United States. EFSA confirmed safety evidence and set adult UL to 7.0 mg/day (lower in children)[41] For the purpose of displaying U.S. food and dietary supplements, the amount of vitamins or minerals during serving is expressed as a percentage of the day value (%DV). There is information to set the appropriate intake, but fluoride does not have a daily value and does not need to be displayed on the food label. [42] Daily estimated intake Of fluoride may vary greatly depending on the source of exposure. Values in the range from 0.46 to 3.6-5.4 mg/day have been reported in some(IPCS, 1984) In areas where water is fluoride, fluoride can be expected to be an important source of fluoride, but fluoride is naturally present in virtually all foods and beverages at a wide concentration. The maximum safety day amount of fluoride is 10 mg / day per day in adults (USA) or 7 mg / day (European Union). [39] [41] The upper limit of fluoride intake is set at 0.10 mg/kg/day for all sources (fluorine water, food, beverages, fluoride dental products and nutritional fluoride supplements), infants, and children up to 8 years of age. [44] Fluoride-containing food/beverage fluoride (1000 g/ppm per mg) fluoride (mg per part) tea (brewing) 3.73 1 cup, 240 g (8 ounces) 0.884 raisins, seedless 2.34 small boxes, 43 g (1.5 ounces) 0.101 table wine 1.53 bottles, 750 ml (26.4 fl oz) 1.150 municipal tap water, (fluorination) 0.81 recommended daily intake, 3 liters (0.79 US gal) 2.433 baked potatoes, Russet 0.455 140 g (0.3 lb) 0.078 ram 0.32 chops, 170 g (6 ounces) 0.054 carrots 0.03 1 large carrot, 72 g (2.5 ounces) 0.002 Source: Data obtained from the U.S. Department of Agriculture, National Nutrition Database[45] According to the main article on safety The highest level of nutrient intake per day that is likely to have no risk of adverse effects on health specifies 10 mg/day for most people corresponding to 10 L of water fluoride without risk. For infants and infants, the water and food sources of fluoride from 0.7 mg/d to 2.2 mg/d.[46] in infants include local water fluoride, seafood, tea, and gelatin. Soluble fluoride salts with the highest sodium fluoride were toxic, resulting in both accidental and self-lying deaths from acute intoxication. The lethal dose of most adults is estimated to be 5 to 10 g (equivalent to elemental fluoride / kg body weight of 32 to 64 mg / kg). [48] [49] Cases of fatal poisoning in adults, including 4 grams of sodium fluoride, have been recorded [51], and a dose of 120 g of sodium fluoride is alive. [52] For sodium fluorosiate (Na2SiF6), the median lethal dose in rats (LD50) is 0.125 g/kg, which is equivalent to 12.5 g for 100 kg adults. [53] Treatment involves oral administration of calcium hydroxide or calcium chloride to prevent further absorption, and with the injection of calcium gluconate to increase the calcium concentration in the blood. Hydrogen fluoride is more dangerous than salts such as NaF because it is corrosive and volatile, and can cause fatal exposure when inhaled or in contact with the skin.Gluconate gel is a normal antidote. At high doses used in the treatment of osteoporosis, sodium fluoride can cause leg pain and incomplete stress fractures when the dose is too high. It also stimulates the stomach, sometimes severe enough to cause ulcers.Slow release of sodium fluoride and intestinal coating versions do not have stomach side effects in an important way and have mild and less frequent complications of bone. At low doses used for water fluorination, only fluorosis of teeth, which can change the appearance of teeth during tooth development, is an obvious side effect. This is almost mild and is unlikely to represent the actual impact on aesthetic appearance or public health. Fluoride was known to enhance the measurement of bone mineral density in the lumbar spine, but it had no effect on vertebral fractures and caused more non-vertebral fractures. Common urban myths claim that the Nazis used fluoride in concentration camps, but there is no historical evidence to prove this claim. In areas where high concentrations of fluoride are naturally present in groundwater used in drinking water, both dental and skeletal fluorescence are prevalent and likely to be severe [hazard maps of fluoride in groundwater show that about one-third of the human population drinks water from groundwater resources.] Of these, about 10 percent and about 300 million people get water from groundwater resources contaminated with aerenic or fluoride. These trace elements are mainly derived from minerals. [61] The map is available for potential problem well locations. [62] Locally concentrated fluoride solution is corrosive. Wear nitrile rubber gloves when handling fluoride compounds. The risk of a solution of fluoride salt depends on the concentration. In the presence of strong acids, fluoride salts release hydrogen fluoride, which is particularly corrosive towards glass. [4] Other derivatives of organic and inorganic anions are produced from fluoride, used as an etchant for glass [64] tetrafluoroholophosphate tetrafluorophosphate, hexafluorophosphate used in organic metal synthesis as an electrolyte of commercially available secondary batteries. For more information on trifluoromethane sulfonic acid, see Medical Portal Fluorine-19 Nuclear Magnetic Resonance Spectroscopy Fluoride Deficiency Fluoride Selective Electrode Fluoride Treatment Fluoride Sodium Fluoride Monofluorophosphate . Pubchem Project. U.S.A.: National Biotechnology Information Center Identification. Chase, M.W. (1998). Fluorine Anion NIST: 1-1951.Acquired on July 4, 2012. The Citation Journals require journals (Help) and Wells, J.C. (2008). Longman Pronunciation Dictionary (3rd) Harlow, England: Pearson Education Limited/Longman. p. 313.According to this source, /ˈfluːəraɪd/ is a possible pronunciation in British English. 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