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This article is about Al(OH)3. For AlOH, see aluminium monohydroxide. Aluminium hydroxide Unit cell ball and stick model of aluminium hydroxide Sample of aluminium hydroxide in a vial Names Preferred IUPAC name Aluminium hydroxide Systematic IUPAC name Trihydroxidoaluminium Other names Aluminic acidAluminic hydroxideAlumanetriolAluminium(III) hydroxideAluminium hydroxideAluminium trihydroxideHydrated aluminaOrthoaluminic acid Identifiers CAS Number 21645-51-2 Y 3D model (JSmol) Interactive image ChEBI CHEBI:33130 Y ChEMBL ChEMBL1200706 N ChemSpider 8351587 Y DrugBank DB06723 ECHA InfoCard 100.040.433 KEGG D02416 PubChem CID 10176082 RTECS number BD0940000 UNII 5QB0T2IUN0 Y CompTox Dashboard (EPA) DTXSID2036405 InChI InChI=1S/Al.3H2O/h;3*1H2/q+3;;;/p-3 YKey: WNROPYMDJYEPJX-UHFFFAOYSA-K InChI=1/Al.3H2O/h;3*1H2/q+3;;;/p-3Key: WNROPYMDJYEPJX-DFZHHIFOAJ SMILES [OH-].[OH-].[OH-].[Al+3] Properties[1][2] Chemical formula Al(OH)3 Molar mass 78.003 g·mol−1 Appearance White amorphous powder Density 2.42 g/cm3, solid Melting point 300 °C (572 °F; 573 K) Solubility in water 0.0001 g/(100 mL) Solubility product (Ksp) 3×10−34 Solubility soluble in acids and alkalis Acidity (pKa) >7 Isoelectric point 7.7 Thermochemistry[3] Std enthalpy offormation (ΔfH298) −1277 kJ·mol−1 Pharmacology[4] ATC code A02AB01 (WHO) Hazards GHS labelling: Pictograms no GHS pictograms Hazard statements no hazard statements Precautionary statements P261, P264, P271, P280, P304+P340, P305+P351+P338, P312, P337+P313 NFPA 704 (fire diamond) 1 0 0 Flash point Non-flammable Lethal dose or concentration (LD, LC): LD50 (median dose) >5000 mg/kg (rat, oral) Safety data sheet (SDS) External MSDS Related compounds Other anions None Related Cations None Related compounds Aluminium oxideAluminium(III) hydroxideIndium(III) hydroxideThallium(III) hydroxideScandium(III) hydroxideSodium oxideAluminium oxide hydroxide Except where otherwise noted, data are given for materials in their standard state (at 25 °C [77 °F], 100 kPa). N verify (what is YN ?) Infobox references Chemical compound Aluminium hydroxide, Al(OH)3, is found as the mineral gibbsite (also known as hydrargillite) and its three much rarer polymorphs: bayerite, doyleite, and nordstrandite. Aluminium hydroxide is amphoteric, i.e., it has both basic and acidic properties. Closely related are aluminium oxide hydroxide, AlO(OH), and aluminium oxide or alumina (Al2O3), the latter of which is also amphoteric. These compounds together are the major components of the aluminium ore bauxite. Aluminium hydroxide also forms a gelatinous precipitate in water. Al(OH)3 is built of double layers of hydroxyl groups with aluminium ions occupying two-thirds of the octahedral holes between the two layers.[5][6] Four polymorphs are recognized.[7] All feature layers of octahedral aluminium hydroxide units, with hydrogen bonds between the layers. The polymorphs differ in terms of the stacking of the layers. All forms of Al(OH)3 crystals are hexagonal [disputed – discuss]: gibbsite is known as γ-Al(OH)3 [8] or α-Al(OH)3 [citation needed] bayerite is also known as α-Al(OH)3[8] or β-alumina trihydrate[citation needed] nordstrandite is also known as Al(OH)3[8] doyleite Hydrargillite, once thought to be aluminium phosphate. Nonetheless, both gibbsite and hydrargillite refer to the same polymorphism of aluminium hydroxide, with gibbsite used most commonly in the United States and hydrargillite used more often in Europe. Hydrargillite is named after the Greek words for water (hydr[o-]) and clay (argillos). Aluminium hydroxide is amphoteric. In acid, it acts as a Brønsted–Lowry base. It neutralizes the acid, yielding a salt:[9] 3 HCl + Al(OH)3 → AlCl3 + 3 H2O In bases, it acts as a Lewis acid by binding hydroxide ions:[9] Al(OH)3 + OH− → [Al(OH)4]− Red mud reservoirs (this one in Stade, Germany) contain the corrosive residues from the production of aluminium hydroxide. Virtually all the aluminium hydroxide used commercially is manufactured by the Bayer process[10] which involves dissolving bauxite in sodium hydroxide at temperatures up to 270 °C (518 °F). The waste solid, bauxite tailings, is removed and aluminium hydroxide is precipitated from the remaining solution of sodium aluminate. This aluminium hydroxide can be converted to aluminium oxide or alumina by calcination.[citation needed] The residue or bauxite tailings, which is mostly iron oxide, is highly caustic due to residual sodium hydroxide. It was historically stored in lagoons; this led to the Ajka alumina plant accident in 2010 in Hungary, where a dam bursting led to the drowning of nine people. An additional 122 sought treatment for chemical burns. The mud contaminated 40 square kilometres (15 sq mi) of land and reached the Danube. While the mud was considered non-toxic due to low levels of heavy metals, the associated slurry had a pH of 13.[11] Aluminium hydroxide finds use as a fire retardant filler for polymer applications. It is selected for these applications because it is colorless (like most polymers), inexpensive, and has good fire retardant properties.[12] Magnesium hydroxide and mixtures of huntite and hydromagnesite are used similarly.[13][14][15][16][17] These mixtures start to decompose at temperatures around 180 °C (356 °F) to 220 °C (428 °F) (depending on the type of aluminium hydroxide used), absorbing a considerable amount of heat in the process and giving off water vapour. The decomposition rate of aluminium hydroxide increases with an increase in temperature, with a reported maximum rate at 250 °C (482 °F).[18] In addition to behaving as a fire retardant, it is very effective as a smoke suppressant in a wide range of polymers, most especially in polyesters, acrylics, ethylene vinyl acetate, epoxies, polyvinyl chloride (PVC), rubber,[19] as well as in wood-based products.[20][21][22] Aluminium hydroxide is used as filler in some artificial stone compound material, often in acrylic resin.[citation needed] Aluminium hydroxide is a feedstock for the manufacture of other aluminium compounds: calcined aluminas, aluminium sulfate, polyaluminium chloride, aluminium chloride, zeolites, sodium aluminate, activated alumina, and aluminium nitrate.[6] Freshly precipitated aluminium hydroxide forms gels, which are the basis for the application of aluminium salts as flocculants in water purification. This gel crystallizes with time. Aluminium hydroxide gels can be dehydrated (e.g. using water-miscible non-aqueous solvents like ethanol) to form an amorphous aluminium hydroxide powder, which is readily soluble in acids. Heating converts it to activated aluminas, which are used as desiccants, adsorbent in gas purification, and catalyst supports.[12] Under the generic name "algeldrate", aluminium hydroxide is used as an antacid in humans and animals (mainly cats and dogs). It is preferred over other alternatives such as sodium bicarbonate because Al(OH)3, being insoluble, does not increase the pH of stomach above 7, and hence does not trigger secretion of excess acid by the stomach. Brand names include Alu-Cap, Aludrox, Gaviscon or Pepsamar. It reacts with excess acid in the stomach, reducing the acidity of the stomach content,[23][24] which may relieve the symptoms of ulcers, heartburn or dyspepsia. Such products can cause constipation, because the aluminium ions inhibit the contractions of smooth muscle cells in the gastrointestinal tract, slowing peristalsis and lengthening the time needed for stool to pass through the colon.[25] Some such products are formulated to minimize such effects through the inclusion of equal concentrations of magnesium hydroxide or magnesium carbonate, which have counterbalancing laxative effects.[26] This compound is also used to control hyperphosphatemia (elevated phosphate, or phosphorus, levels in the blood) in people and animals suffering from kidney failure. Normally, the kidneys filter excess phosphate from the blood, but kidney failure can cause phosphate to accumulate. The aluminium salt, when ingested, binds to phosphate in the intestines and reduce the amount of phosphorus that can be absorbed.[27][28] Precipitated aluminium hydroxide is included as an adjuvant in some vaccines (e.g. anthrax vaccine). One of the well-known brands of aluminium hydroxide adjuvant is Alhydrogel, made by Brenntag Biosector.[29][full citation needed][dead link] Since it absorbs protein well, it also functions to stabilize vaccines by preventing the proteins in the vaccine from precipitating or sticking to the walls of the container during storage. Aluminium hydroxide is sometimes called "alum", a term generally reserved for one of several sulfates.[citation needed] Vaccine formulations containing aluminium hydroxide stimulate the immune system by inducing the release of uric acid, an immunological danger signal. This strongly attracts certain types of monocytes which differentiate into dendritic cells. The dendritic cells pick up the antigen, carry it to lymph nodes, and stimulate T cells and B cells.[30] It appears to contribute to induction of a good Th2 response, so is useful for immunizing against pathogens that are blocked by antibodies. However, it has little capacity to stimulate cellular (Th1) immune responses, important for protection against many pathogens,[31] nor is it useful when the antigen is peptide-based.[32] In the 1960s and 1970s it was speculated that aluminium was related to various neurological disorders, including Alzheimer's disease.[33][34] Since then, multiple epidemiological studies have found no connection between exposure to environmental or swallowed aluminium and neurological disorders, though injected aluminium was not looked at in these studies.[35][36][37] Neural disorders were found in experiments on mice motivated by Gulf War illness (GWI). Aluminium hydroxide injected in doses equivalent to those administered to the United States military, showed increased reactive astrocytes, increased apoptosis of motor neurons and microglial proliferation within the spinal cord and cortex.[38] ^ For solubility product: "Solubility product constants". Archived from the original on 15 June 2012. 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There is another way to use Hess' Law. It does not use the full chemical equations and it is usually presented like this: ΔH rxno = Σ ΔH fo (products) − Σ ΔH fo (reactants) Here's another to write this form of Hess' Law, one that slightly varies from the above manner: ΔH rxno = Σ ΔH f, productso − Σ ΔH f, reactantso The "rxn" above is a common way to abbreviate "reaction." All it means is that we are discussing the enthalpy of a generic reaction, not any specific one. I'll explain the above equation using an example problem. Example #1: Calculate the standard enthalpy of combustion for the following reaction: C2H5OH(ℓ) + 7/2O2(g) ---> 2CO2(g) + 3H2O(ℓ) Before launching into the solution, notice I used "standard enthalpy of combustion." This is a very common chemical reaction, to take something and combust (burn) it in oxygen. It is so common that the phrase "standard enthalpy of combustion" is used alot and gets its own symbol: ΔH°comb. The key to solving this problem is to have a table of standard enthalpies of formation handy. In case you missed it, look at the equation up near the top and see the subscripted f. What we are going to do is sum up all the product enthalpies of formation and then subtract the summed up reactant enthalpies of formation. Also, we need to have the equation balanced, so be sure to remember to check for that. Fractional coefficients are OK. Like this: ΔH combo = [(−393.5) + 3 (−286)] − [(−278) + (7/2) (0)] The boldfaced values are the coefficients and the other ones are the standard enthalpy of formation for the four substances involved. Since oxygen is an element in its standard state, its enthalpy of formation is zero. Doing the math gives us ΔH combo = −1367 kJ/mol of ethyl alcohol. The ΔH fo values (−393.5, −286, −278 and zero) were looked up in a reference source. Textbooks which teach this topic will have an appendix of the values. Make sure you find it and figure out how to use it. If you are not too clear on what the term "standard enthalpy of formation" means, please look here. Example #2: Calculate the standard enthalpy of formation of glucose: C6H12O6(s) + 6O2(g) ---> 6CO2(g) + 6H2O(ℓ) To solve this problem, we must know the following ΔH fo values: −1275.0 O2(gzero CO2(g)−393.5 H2O(ℓ)−285.8 All the above values have units of kJ/mol because these are standard values. All standard enthalpies have the unit kJ/mol. As a brief reminder, here is the chemical reaction for the standard enthalpy of glucose: 6C(s, graphite) + 6H2(g) + 3O2(g) ---> C6H12O6(s) Each standard enthalpy value is associated with a chemical reaction. The reaction will always form one mole of the target substance (glucose in the example) in its standard state. The target substance is always formed from elements in their respective standard states. Note how the standard state for carbon is graphite, not diamond or buckerministerfullerene. Remember also that all elements in their standard state have an enthalpy of formation equal to zero. The solution ΔH combo = [6 (−393.5) + 6 (−285.8)] − [(1) (−1275) + (6) (0)] The boldfaced values are the coefficients and the other ones are the standard enthalpy of formation for the four substances involved. Since oxygen is an element in its standard state, its enthalpy of formation is zero. Doing the math gives us ΔH combo = −2801 kJ/mol of glucose. Example #3: Calculate the standard enthalpy of glucose, given the following values: ΔH comb, glucoseo = −2800.8 kJ/mol ΔH f, CO2o = −393.5 kJ/mol ΔH f, H2Oo = −285.8 kJ/mol Solution: −2800.8 = [(6) (−393.5) + (6) (−285.8)] − [(ΔH f, glucoseo) + (6) (0)] Do you see what I did? All the enthalpies of formation on the right-hand side and the ΔH comb on the left-hand side. By the way, this is a common test question. Be prepared. Here's what happened: 1) First of all, this is the reaction we want answer for: 6C(s, graphite) + 6O2(g) ---> C6H12O6(s) + 6O2(g) ---> 6CO2(g) + 6H2O(ℓ) 2) We see the reactions to be added, in the manner of Hess' Law: C6H12O6(s) + 6O2(g) ---> 6CO2(g) + 6H2O(ℓ) 1/2O2(g) --->H2O(ℓ) Flip the first reaction and multiply the other two by six. Then add the three reactions together. If you do it right, you should recover the reaction mentioned just above in (1). Example #4: Complete combustion of 1.00 mol of acetone (C3H6O) liberates 1790 kJ: C3H6O(ℓ) + 4O2(g) ---> 3CO2(g) + 3H2O(ℓ) ΔH combo = −1790 kJ Using this information together with the data below (values in kJ/mol), calculate the ΔH fo for acetone. C3H6O = 0 ΔH f, CO2o = −393.5 kJ ΔH f, producto = Σ ΔH f, reactantso 2) Sustitute values into equation: −1790 = [(3) (−393.5) + (3) (−285.8)] − [(ΔH f, acetoneo) + (4) (0)] −1790 = −2037.9 − ΔH°f, acetone 247.99 = ΔH f, acetoneo ΔH f, acetoneo = −247.99 kJ/mol To three sig figs, the value is −248 kJ/mol. Example #5: The standard enthalpy of formation of hexane can be determined indirectly. Calculate the standard enthalpy of formation of hexane using the enthalpies of combustion of C6H14(ℓ) + 19/2O2(g) ---> 6CO2(g) + 7H2O(ℓ) ΔH = −4163.0 kJ C(s, gr) + O2(g) ---> CO2(g) ΔH = −393.5 kJ H2(g) + 1/2O2(g) ---> H2O(ℓ) ΔH = −285.5 kJ 2) To obtain the target reaction (the production of hexane), we must do the following: a) reverse the first equation b) multiply the second equation by 6 c) multiply the third equation by 7 3) After doing that, we obtain: C6H14(ℓ) + 19/2O2(g) ---> 6CO2(g) + 7H2O(ℓ) ΔH o = +4163.0 kJ 6C(s, gr) + 6O2(g) ---> 6CO2(g) ΔH o = −2361.0 kJ 7H2(g) + 7/2O2(g) ---> 7H2O(ℓ) ΔH o = −2000.6 kJ By the way, the second equation (presented as the enthalpy of combustion of carbon) is also the equation for the formation of carbon dioxide. The third equation (presented as the combustion of hydrogen gas) is also the formation equation for water in its standard state (liquid). The moral of the story? Sometimes terms overlap. The −393.5 value is the enthalpy of combustion of carbon. It is also the formation enthalpy for carbon dioxide. And the enthalpies to obtain the enthalpies of combustion for CO2(g) and H2O(ℓ) are also their enthalpies of formation. By the way, this is a common test question. Be prepared. Here's what happened: 1) First of all, this is the reaction we want answer for: 6C(s, graphite) Example #6: Ammonia reacts with oxygen to form nitrogen dioxide and steam, as follows: 4NH3(g) + 7O2(g) ---> 4NO2(g) + 6H2O(g) Using the following standard enthalpies of formation (given in kJ/mol), calculate the ΔH rxno. NH3(g)−45.90 N2(g)33.1 H2O(ℓ)−241.83 Note that water is given as a gas. The usual problem of this type uses water as a liquid. Not in this one. Solution: Use Hess' Law: ΔH rxno = [(4) (+33.1) + (6) (−241.83)] − [(4) (−45.90) + (7) (0)] ΔH rxno = −1134.98 kJ = −1135 kJ Note that the units kJ/mol are NOT used. Example #7: The standard enthalpy change, ΔH°, for the thermal decomposition of silver nitrate according to the following equation is +78.67 kJ: AgNO3(s) ---> AgNO2(s) + 1/2O2(g) The standard enthalpy of formation of AgNO3(s) is −123.02 kJ/mol. Calculate the standard enthalpy of formation of AgNO2(s) Solution: 1) Let's write what we know: AgNO3(s) ---> AgNO2(s) + 1/2O2(g) ΔH o = +78.67 kJ Ag(s) + 1/2N2(g) + 3/2O2(g) ---> AgNO3(s) ΔH fo = −123.02 kJ) = −44.35 kJ (this is the answer) When the two equations are added together, the AgNO3 cancels out as does 1/2O2(g) and we are left with the formation equation for AgNO2(s), the equation given in step 2. Note that the two data equations did not require any modification (flipping or multiplying by a factor). Example #8: Using standard enthalpies of formation, calculate the heat of combustion per mole of gaseous water formed during the complete combustion of ethane gas. Solution: 1) The combustion reaction for ethane is: 2C2H6 + 7O2 ---> 4CO2 + 6H2O 2) The balanced equation for the combustion of C2H6 (ethane) is: 2C2H6 + 7O2 ---> 4CO2 + 6H2O 2) The balanced equation for the combustion of ethylene glycol.