Chapter 1 : Group 1: Properties of Alkali Metals - Chemistry LibreTexts

Group 1 elements: Trend of change in the physical properties: The atomic radius (atomic size) of alkali metals increases gradually. down the group.. Reason: The number of shells occupied with electrons increases down the group.

The low melting points of the alkali metals are a direct result of the large interatomic distances in their crystals and the weak bond energies associated with such loose arrays. These same factors are responsible for the low densities, low heats of fusion, and small changes in volume upon fusion of the metals. Lithium, sodium , and potassium are less dense than water. The large size of an alkali metal atom and the resulting low density of the metal results from the presence of only one, weakly bound electron in the large outer s-type orbital. Upon going from the noble-gas configuration of argon atomic number 18 to potassium atomic number 19 , the added electron goes into the large 4s orbital rather than the smaller 3p orbital. When, however, potassium, rubidium , or cesium metals are subjected to increasing pressure up to one-half million atmospheres or more , a number of phase transitions occur. Ultimately, occupation of a d-type orbital becomes preferred over that of the s-orbital, with the result that these alkali metals resemble transition metals. Under such circumstances, alloys with transition metals such as iron can form, a result that does not occur at low pressures. The alkali metals have played an important role in quantum physics. Some alkali metal isotopes, such as rubidium, are bosons. In this state, the cluster of atoms is in a single quantum state and exhibits macroscopic behaviours normally seen only with atomic-sized particles. Chemical properties Since the alkali metals are the most electropositive the least electronegative of elements, they react with a great variety of nonmetals. In its chemical reactivity, lithium more closely resembles Group 2 IIa of the periodic table than it does the other metals of its own group. It is less reactive than the other alkali metals with water , oxygen , and halogens and more reactive with nitrogen , carbon , and hydrogen. Therefore, neutral compounds with oxygen can be readily classified according to the nature of the oxygen species involved. Ionic oxygen species include the oxide, O^{2-}, peroxide , O, superoxide, O^{2-}, and ozonide O Rubidium and cesium and, possibly, potassium also form the sesquioxide, M_4O_6, which contains two peroxide anions and one superoxide anion per formula unit. Lithium forms only the monoxide and the peroxide. All the alkali metals react directly with oxygen; lithium and sodium form monoxides, Li_2O and Na_2O, and the heavier alkali metals form superoxides, MO_2. The rate of reaction with oxygen, or with air, depends upon whether the metals are in the solid or liquid state , as well as upon the degree of mixing of the metals with the oxygen or air. In the liquid state, alkali metals can be ignited in air with ease, generating copious quantities of heat and a dense choking smoke of the oxide. The close approach of the small lithium ion to the oxygen atom results in the unusually high free energy of formation of the oxide. The peroxides Li_2O_2 and Na_2O_2 can be made by passing oxygen through a liquid- ammonia solution of the alkali metal, although sodium peroxide is made commercially by oxidation of sodium monoxide with oxygen. Sodium superoxide NaO_2 can be prepared with high oxygen pressures, whereas the superoxides of rubidium, potassium, and cesium can be prepared directly by combustion in air. By contrast, no superoxides have been isolated in pure form in the case of lithium or the alkaline-earth metals, although the heavier members of that group can be oxidized to the peroxide state. The cyanides of potassium, rubidium, and cesium, which are less stable than the lower oxides, can be prepared by the reaction of the superoxides with ozone. The rate of the reaction depends on the degree of metal surface presented to the liquid. With small metal droplets or thin films of alkali metal, the reaction can be explosive. The rate of the reaction of water with the alkali metals increases with increasing atomic weight of the metal. With the heavier alkali metals, the hydroxides are highly soluble; thus, they are removed readily from the reacting surface, and the reaction can proceed with unabated vigour. The reaction involves equimolar mixtures that is, equal numbers of atoms or molecules of the alkali metal and water to form a mole an amount equal to that of the reactants of alkali metal hydroxide and half a mole of hydrogen gas. These reactions are highly exothermic give off heat , and the hydrogen that is generated can react with oxygen to increase further the heat that is generated. Reactions with nonmetals Of the alkali metals, only lithium reacts with nitrogen, and it forms a nitride Li_3N. In this respect it is more similar to the alkaline-earth metals than to the Group 1 metals. Lithium also forms a relatively stable hydride , whereas the

other alkali metals form hydrides that are more reactive. Lithium forms a carbide Li_2C_2 similar to that of calcium. The other alkali metals do not form stable carbides, although they do react with the graphite form of carbon to give intercalation compounds substances in which the metal atoms are inserted between layers of carbon atoms in the graphite structure. The alkali metals can be burned in atmospheres of the various halogens to form the corresponding halides. The alkali metals react with nonmetals in Groups 15 and 16 Va and VIa of the periodic table. Sulfides can be formed by the direct reaction of the alkali metals with elemental sulfur , furnishing a variety of sulfides. Phosphorus combines with the alkali metals to form phosphides with the general formula M_3P. Formation of alloys The characteristics of alloy behaviour in alkali metals can be evaluated in terms of the similarity of the elements participating in the alloy. Elements with similar atomic volumes form solid solutions that is, mix completely in all proportions ; some dissimilarity in atomic volumes results in eutectic-type systems solutions formed over limited concentration ranges , and further dissimilarity results in totally immiscible systems. The high-pressure transition in potassium, rubidium, and cesium that converts these s-type metals to more transition metal-like d-type metals yields atomic volumes that are similar to those of many transition metals at the same pressure. This permits alloys or compounds to form between these alkali metals and such transition metals as nickel or iron. The elements potassium, rubidium, and cesium, which have rather similar atomic volumes and ionization energies, form complete solid solutions and mixed crystals. Sodium, which is a significantly smaller atom than potassium and has a higher ionization energy, tends to form eutectic systems with potassium, rubidium, and cesium. Even greater dissimilarity exists in the atomic volumes of sodium and lithium, resulting in insolubilities of the liquid phases. The consolute temperature the temperature at which the two liquids become completely miscible increases on going from the lithium-sodium alloy system to the lithium-cesium system. There is only one example of solid miscibility in alkaliâ€"alkaline-earth-metal binariesâ€"the lithium-magnesium system, in which the two elements are very similar. Sodium forms compounds only with barium in the alkaline-earth-metal series. The heavier alkali metals all tend to form immiscible liquid phases with the alkaline earth metals. Several elements in Group 12 IIb of the periodic table zinc , cadmium , and mercury react with the alkali metals to form compounds. Mercury forms at least six compounds, commonly termed amalgams, with each of the five alkali metals, and with the exception of the amalgam with lithium, the highest melting point compound in each series has the formula MHg_2. Lithium and sodium also form compounds with cadmium and zinc. Formation of complexes Until the late s there were few complexes of the alkali metal cations with organic molecules. The only significant difference between one alkali cation and another is the size. The synthesis of crown ethers by American chemist Charles J. Thus, the selectivity of a crown ether for a particular cation depends on the ring size. Common crown ethers are crown-4, crown-5, and crown Three crown ethers, each with a complexed alkali metal cation. Even greater affinity for alkali cations was achieved by the synthesis of cryptands by French chemist Jean-Marie Lehn in and spherands by American chemist Donald Cram in These are three-dimensional molecules with an internal cavity or crypt that can completely encapsulate the alkali cation. It should be noted, however, that these molecules are not rigid and that flexibility of the framework can alter the cavity size to accommodate alkali cations of different sizes, although with differences in the strength of complexation. Since the initial syntheses of crown ethers and cryptands, thousands of complexants for cations of various sizes, charges, and geometries have been synthesized. This has led to an entirely new branch of chemistry called supramolecular chemistry. Analytical chemistry of the alkali metals Classical methods of separation and analysis of alkali metals are rather difficult and time consuming. For lithium they include such procedures as selective extraction of lithium chloride into organic solvents and the detection of lithium with azo dyes that give highly sensitive colour reactions in alkaline solutions. A modification of the uranyl acetate test the precipitation of an insoluble sodium salt with uranyl acetate has been used as a standard test for the presence of sodium. The use of a cobaltinitrite solution permits separation of potassium from sodium by precipitation of the insoluble potassium salt. There are essentially no satisfactory analytical methods for rubidium and cesium based on the use of reagents in solution. Classical methods of separation of the alkali metals have been largely supplanted by chromatographic elution. Strongly acidic cation -exchange resins and aqueous acidic solutions are used. Ion-exchange resins that are specific for lithium have been developed.

Macrocyclic compounds such as crown ethers and cryptands that are selective for particular alkali metal ions have been synthesized. The characteristic flame colours of the alkali metals red, yellow, violet, red, and blue for Li, Na, K, Rb, and Cs, respectively are qualitative indicators of the modern analytical methods used to determine the concentrations of alkali-metal salts in aqueous solution. The intensities of the characteristic spectral lines in emission after excitation by a flame or ICP inductively coupled plasma give quantitative measures of the individual alkali metal concentration in the parts per million range or lower. Determination of one alkali metal in the presence of another, however, can result in interference, which can be reduced by using specially prepared standard solutions that contain known amounts of the interfering metals. The analysis of the alkali-metal samples for the presence of nonmetallic elements, such as oxygen , carbon , hydrogen , and nitrogen , requires specialized techniques. The oxygen content of sodium and potassium samples can be determined by extraction of the free alkali metal with mercury , leaving behind mercury-insoluble oxides and carbonates , which can subsequently be analyzed by means of solution methods. The oxygen content of rubidium and cesium can be accurately determined by precise measurement of the freezing point of these two elements. The carbon content of alkali metals can be analyzed by oxidation of the alkali metal in pure oxygen, followed by infrared measurement of the carbon dioxide generated during combustion. For the analysis of nitride in lithium, the nitride commonly is converted to ammonia , and the ammonia is measured by colorimetric analysis.

atomic and physical properties of the group 1 elements This page explores the trends in some atomic and physical properties of the Group 1 elements - lithium, sodium, potassium, rubidium and caesium. You will find separate sections below covering the trends in atomic radius, first ionisation energy, electronegativity, melting and boiling points.

They are silvery-white but we do not usually see them glistening due to a thin layer of alkali metal oxides coating on the surface of the metals. All of them have very low melting points, the melting points drop down the periodic table. Second is the chemical properties of alkali metals. All alkali metals have the electron configuration of ns1, this means that all alkali metals tend to loss one electron to achieve a noble gas electron configuration. Thus, they are highly reactive and are all strong reducing agents, increasingly so down the periodic table. They are able to react with almost all non-metals vigorously. All alkali metals react with water to form strong alkalis and hydrogen: The reaction becomes more violent down the periodic table. Lithium fizzes when reacting with alkali. Sodium and potassium catch fire in water. Rubidium and Caesium explode in water. Lithium reacts violently with oxygen to form the only oxide, lithium oxide, Li_2O. Sodium combusts in air to form sodium peroxide, Na_2O_2. Exposed to air at r. Rubidium and caesium even possess ozonides such as CsO_3. All metals quickly tarnish in air. Caesium is so reactive that burns in air in minutes. Alkali metals are reducing enough to transfer an electron to hydrogen atoms, forming hydride anions. They react with hydrogen to give stable hydrides. Some alkali earth metals also possess this property. Lithium is somehow a bit different from other alkali metals, it has similar chemical properties as magnesium in accordance to the diagonal similarity in periodic table. For instance, lithium nitrate decomposes differently from other alkali metal nitrates, its decomposition releases nitrogen dioxide.

Chapter 3 : Physical Properties Of Group 2A Elements | TutorsOnNet

This page discusses the trends in some atomic and physical properties of the Group 1 elements - lithium, sodium, potassium, rubidium and cesium. Sections below cover the trends in atomic radius, first ionization energy, electronegativity, melting and boiling points, and density. In each element, the.

You can help by adding to it. November Of the six alkaline earth metals, beryllium, calcium, barium, and radium have at least one naturally occurring radioisotope ; magnesium and strontium do not. Beryllium-7 , beryllium , and calcium are trace radioisotopes ; calcium and barium have very long half-lives and thus occur naturally on earth; and all isotopes of radium are radioactive. Calcium is the lightest nuclide to undergo double beta decay. These oxides are basic alkaline when combined with water. The realization that these earths were not elements but compounds is attributed to the chemist Antoine Lavoisier. Later, he suggested that the alkaline earths might be metal oxides, but admitted that this was mere conjecture. Discovery[edit] The calcium compounds calcite and lime have been known and used since prehistoric times. The magnesium compound magnesium sulfate was first discovered in by a farmer at Epsom in England. Strontium carbonate was discovered in minerals in the Scottish village of Strontian in The last element is the least abundant: Beryl , a mineral that contains beryllium, has been known since the time of the Ptolemaic Kingdom in Egypt. Calcium sulfate has been known to be able to set broken bones since the tenth century. Strontium[edit] In , physician Adair Crawford , who had been working with barium, realized that Strontian ores showed different properties than other supposed ores of barium. Strontium was eventually isolated in by Sir Humphry Davy by electrolysis of a mixture of strontium chloride and mercuric oxide. The discovery was announced by Davy on 30 June at a lecture to the Royal Society. Barite , a mineral containing barium, was first recognized as containing a new element in by Carl Scheele , although he was able to isolate only barium oxide. Barium oxide was isolated again two years later by Johan Gottlieb Gahn. Later in the 18th century, William Withering noticed a heavy mineral in the Cumberland lead mines, which are now known to contain barium. Barium itself was finally isolated in when Sir Humphry Davy used electrolysis with molten salts, and Davy named the element barium, after baryta. Later, Robert Bunsen and Augustus Matthiessen isolated pure barium by electrolysis of a mixture of barium chloride and ammonium chloride. The material behaved somewhat similarly to barium compounds , although some properties, such as the color of the flame test and spectral lines, were much different. They announced the discovery of a new element on 26 December to the French Academy of Sciences. Beryllium is one of the rarest elements in seawater, even rarer than elements such as scandium , with a concentration of 0. None of the alkaline earth metals are found in their elemental state, but magnesium and calcium are found in many rocks and minerals: Most strontium is found in the minerals celestite and strontianite.

Chapter 4 : Alkaline earth metal - Wikipedia

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Here is a discussion on physical properties of group IIA elements. Elements included in this group include the beryllium, magnesium, calcium, strontium, barium and radium. Among all the elements, radium is the radioactive element. So, group IIA elements are also termed as alkaline earth metals. These elements have two electrons in their outermost orbital. They have a silvery luster. But, it soon disappears upon exposure to air. They are malleable and ductile but very less when compared to alkali metals. Atomic Volume and Radius: As we go down the group, the number of electrons is less when compared to the increase in the number of available orbitals. Hence, the atomic radius increases gradually as we move down the group. It should be noted that these elements are smaller when compared to alkali metals. This is because of the presence of two electrons in the outermost shell. Hence, the effect of the nucleus on the outermost shell is comparatively more than the alkali metals which have a single electron. Due to presence of two electrons in the outermost shell they can be more densely packed compared to alkali metals. Melting and Boiling Points: These elements have a higher boiling and melting points. Due to presence of two electrons in the valence shell they remain tightly packed in solid state. Alkaline earth metals have a smaller size and higher nuclear charge as a result of which the valence electrons are held strongly. Hence, more amount of energy is required to remove an electron from the valence electron resulting in high ionization energies. As the size to charge ratio is very high, they are highly electropositive in nature. As they are highly electropositive, they are less electronegative. Electronegativity decreases as we move down the group. Beryllium is highly electronegative due to small size. They are good conductors of heat and electricity. This is due to the presence of a two electrons that can very easily move within the crystal lattice of the elements. When the elements are heated, electrons present in the valence shell are excited to higher energy levels. When the excited electrons return back after losing the energy they emit certain amount light. All Alkaline earth metals except beryllium and magnesium emit various colors depending on the degree of excitation. These two elements due to their smaller size have high ionization energies and high excitation energy. Hence, they are not excited to higher energy states and thus no flame colorization. The two electrons in the valence shell can be very easily removed. Hence, all alkaline earth metals act as strong reducing agents. The reducing nature increases as we move down the group. However, the reducing nature of alkaline earth metals is less than their fellow s-block elements, the alkali metals. Solved problems Which of the following elements does not show flame colorization?

Chapter 5 : Group 1 Metals | gcse-revision, chemistry, periodic-table, groupmetals | Revision World

N-m (Â© Nigel Saunders)Physical properties of group 1 elements 1. Copy the table below into your book and complete it (parts have been done for you already).

Butyllithium C_4H_9Li , which is used in the manufacture of synthetic rubber, is prepared by the reaction of butyl bromide C_4H_9Br with metallic lithium. In many respects lithium also shows similarities to the elements of the alkaline-earth group, especially magnesium , which has similar atomic and ionic radii. This similarity is seen in oxidation properties, the monoxide being normally formed in each case. Reactions of organolithium compounds are also similar to the Grignard reactions of organomagnesium compounds, a standard synthetic procedure in organic chemistry. A number of the lithium compounds have practical applications. Lithium hydride LiH , a gray crystalline solid produced by the direct combination of its constituent elements at elevated temperatures, is a ready source of hydrogen, instantly liberating that gas upon treatment with water. It also is used to produce lithium aluminum hydride $LiAlH_4$, which quickly reduces aldehydes, ketones, and carboxylic esters to alcohols. Lithium hydroxide $LiOH$, commonly obtained by the reaction of lithium carbonate with lime, is used in making lithium salts soaps of stearic and other fatty acids; these soaps are widely used as thickeners in lubricating greases. Lithium hydroxide is also used as an additive in the electrolyte of alkaline storage batteries and as an absorbent for carbon dioxide. Other industrially important compounds include lithium chloride $LiCl$ and lithium bromide $LiBr$. They form concentrated brines capable of absorbing aerial moisture over a wide range of temperatures; these brines are commonly employed in large refrigerating and air-conditioning systems. Lithium fluoride LiF is used chiefly as a fluxing agent in enamels and glasses. Nuclear properties Lithium, which exhibits no natural radioactivity, has two isotopes of mass number 6 Lithium was used in as the target metal in the pioneering work of British physicist John Cockcroft and Irish physicist Ernest Walton in transmuting nuclei by artificially accelerated atomic particles; each lithium nucleus that absorbed a proton became two helium nuclei. The bombardment of lithium-6 with slow neutrons produces helium and tritium $3H$; this reaction is a major source of tritium production. Tritium so produced is employed in the manufacture of hydrogen bombs, among other uses such as providing a radioactive hydrogen isotope for biological research. Lithium has potential value as a heat-transfer fluid for high power-density nuclear reactors. The isotopes lithium-8 half-life 0. Biological properties The widespread occurrence of lithium in plants results in a wide, although low-level, distribution of lithium in animals. Lithium salts have complex effects when absorbed into the body. They are not highly toxic, although high levels can be fatal. The use of lithium salts and mineral water containing them to treat gout unsuccessfully and to ward off depression successfully dates to the last half of the 19th century but fell into medical disrepute in the early 20th century. The use of lithium carbonate to treat manic-depression also known as bipolar disorder was demonstrated clinically in Fears about lithium toxicity delayed its approval for many years, but it is now the major drug for the treatment of manic episodes and for maintenance therapy in bipolar patients.

Chapter 6 : lithium | Definition, Properties, Use, & Facts | calendrierdelascience.com

As a group the alkali metals are the least electronegative of the elements, ranging from to on the scale, while the alkaline earths, the next group on the table, have electronegativities ranging from about to

This page discusses the trends in some atomic and physical properties of the Group 1 elements - lithium, sodium, potassium, rubidium and cesium. Trends in Atomic Radius The chart below shows the increase in atomic radius down the group. The radius of an atom is governed by two factors: The number of layers of electrons around the nucleus The attraction the outer electrons feel from the nucleus Compare the electronic configurations of lithium and sodium: The positive charge on the nucleus is canceled out by the negative charges of the inner electrons. This effect is illustrated in the figure below: This is true for each of the other atoms in Group 1. The only factor affecting the size of the atom is the number of layers of inner electrons which surround the atom. More layers of electrons take up more space, due to electron-electron repulsion. Notice that first ionization energy decreases down the group. Ionization energy is governed by three factors: Down the group, the increase in nuclear charge is exactly offset by the increase in the number of inner electrons. However, the distance between the nucleus and the outer electrons increases down the group; electrons become easier to remove, and the ionization energy falls. Trends in Electronegativity Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons. It is usually measured on the Pauling scale, on which the most electronegative element fluorine is given an electronegativity of 4. A graph showing the electronegativities of the Group 1 elements is shown above. Each of these elements has a very low electronegativity when compared with fluorine, and the electronegativities decrease from lithium to cesium. Picture a bond between a sodium atom and a chlorine atom. This is illustrated in the figure below: The electron pair is so close to the chlorine that an effective electron transfer from the sodium atom to the chlorine atom occurs—the atoms are ionized. This strong attraction from the chlorine nucleus explains why chlorine is much more electronegative than sodium. Now compare this with a lithium-chlorine bond. The net pull from each end of the bond is the same as before, but the lithium atom is smaller than the sodium atom. In some lithium compounds there is often a degree of covalent bonding that is not present in the rest of the group. Lithium iodide, for example, will dissolve in organic solvents; this is a typical property of covalent compounds. This corresponds with a decrease in electronegativity down Group 1. With the exception of some lithium compounds, the Group 1 elements each form compounds that can be considered ionic. Trends in melting and boiling points The figure above shows melting and boiling points of the Group 1 elements. Both the melting and boiling points decrease down the group. When any of the Group 1 metals is melted, the metallic bond is weakened enough for the atoms to move more freely, and is broken completely when the boiling point is reached. The decrease in melting and boiling points reflects the decrease in the strength of each metallic bond. As the atoms increase in size, the distance between the nuclei and these delocalized electrons increases; therefore, attractions fall. The atoms are more easily pulled apart to form a liquid, and then a gas. The increased charge on the nucleus down the group is offset by additional levels of screening electrons. As before, the trend is determined by the distance between the nucleus and the bonding electrons. This trend is shown in the figure below: Fewer sodium atoms than lithium atoms, therefore, can be packed into a given volume. A given number of sodium atoms will weigh more than the same number of lithium atoms. Therefore, 1 cm3 of sodium contains fewer atoms than the same volume of lithium, but each atom weighs more. Mathematical calculations are required to determine the densities.

Physical Properties of Group 15 Elements. Physical Properties of Group 15 Elements. Skip navigation Sign in. Search. Loading Close. This video is unavailable. Watch Queue Queue.

Check new design of our homepage! Well, they belong to the same group in the periodic table, and the members of this group there are 6 elements in all are known as alkali metals. This ScienceStruck article talks about the physical and chemical properties of these elements. ScienceStruck Staff Alkali metals belong to the Group 1 of the periodic table. Being highly reactive in nature, they are mostly found in the form of compounds and not in elemental form. Most of the properties of hydrogen are very different from those of the alkali metals. Only under extremely high pressure, hydrogen shares some properties with the members of this group. Physical Properties Many of the physical properties of alkali metals are very similar to that of other metals. However, there are certain physical properties that make them different from other elements. The densities of alkali metals are much lower when compared to other metals. Unlike other metals, the elements of the alkali metal group are soft substances and can be cut with a knife. In any given period of the periodic table, the atomic radii of the alkali metals is the largest, and the atomic radius increases as one moves down the group. The members of this group have low ionization energies, which is the energy required to remove the outermost electron from the orbit. The ionization energy decreases as one moves down the group. They are good conductors of heat and electricity. They have very low boiling and melting points. Except for cesium, which has a light golden color, alkali metals are silvery white in color. They are lustrous and have a shiny surface. They are highly malleable and ductile, which means they can be easily drawn into thin wires or beaten into thin sheets. They form colored flames when burning in oxygen. Chemical Properties The alkali metals are highly electropositive, which means they readily lose their electrons to form a univalent cation positive ion , and hence, are highly reactive. This is because of the presence of a loosely bound single electron in the outermost shell of these metals. The chemical reactivity of alkali metals increases as we move down the group. They easily react with other elements, especially halogens, to form ionic compounds. Alkali metals also have a strong reducing property. Following are some of the important reactions of alkali metals: Reaction with Oxygen Any alkali metal, on coming in contact with air or oxygen, starts burning and oxides are formed in the process. At the end of the chemical reaction, lithium gives lithium monoxide LiO , sodium gives sodium peroxide Na_2O_2 , and other alkali metals give superoxides that is, each alkali metal atom forms bonds with two oxygen atoms. Oxides of alkali metals are basic in nature and are soluble in water, forming alkali metal hydroxides when dissolved into water. The table below enlists the color of the flame emitted when the alkali metals burn in oxygen.

Chapter 8 : BBC - GCSE Bitesize: Group 1 appearance

Here is a discussion on physical properties of group IIA elements. Elements included in this group include the beryllium, magnesium, calcium, strontium, barium and radium.

However, given that ununennium is only the first period 8 element on the extended periodic table , it may well be discovered in the near future through other reactions, and indeed an attempt to synthesise it is currently ongoing in Japan. Hydrogen and helium are most common, from the Big Bang. The next three elements lithium, beryllium , and boron are rare because they are poorly synthesised in the Big Bang and also in stars. The two general trends in the remaining stellar-produced elements are: Iron is especially common because it represents the minimum energy nuclide that can be made by fusion of helium in supernovae. This rule argues that elements with odd atomic numbers have one unpaired proton and are more likely to capture another, thus increasing their atomic number. In elements with even atomic numbers, protons are paired, with each member of the pair offsetting the spin of the other, enhancing stability. The heavier alkali metals are also less abundant than the lighter ones as the alkali metals from rubidium onward can only be synthesised in supernovae and not in stellar nucleosynthesis. Lithium is also much less abundant than sodium and potassium as it is poorly synthesised in both Big Bang nucleosynthesis and in stars: In turn, the natural history of the Earth caused parts of this planet to have differing concentrations of the elements. The mass of the Earth is approximately 5. It is composed mostly of iron Due to planetary differentiation , the core region is believed to be primarily composed of iron Potassium, rubidium and caesium are also incompatible elements due to their large ionic radii. Potassium makes up approximately 1. Other solid deposits include halite , amphibole , cryolite , nitratine , and zeolite. Commercially, the most important lithium mineral is spodumene , which occurs in large deposits worldwide. It occurs naturally in the minerals leucite , pollucite , carnallite , zinnwaldite , and lepidolite , [57] although none of these contain only rubidium and no other alkali metals. Hence, all the alkali metals are soft and have low densities , [31] melting [31] and boiling points , [31] as well as heats of sublimation , vaporisation , and dissociation. The chemistry of francium is not well established due to its extreme radioactivity ; [31] thus, the presentation of its properties here is limited. What little is known about francium shows that it is very close in behaviour to caesium, as expected. The physical properties of francium are even sketchier because the bulk element has never been observed; hence any data that may be found in the literature are certainly speculative extrapolations.

Chapter 9 : what are the characteristics of group 1 elements? (alkali metals)? | Yahoo Answers

They are called alkali metals because they form strongly alkaline hydroxides with water. In this article, we will talk about the electronic configurations, ionization enthalpy, hydration enthalpy and atomic, ionic radii and other physical and chemical properties of the group 1 alkali metals.

This page explores the trends in some atomic and physical properties of the Group 2 elements: Sections below cover the trends in atomic radius, first ionization energy, electronegativity, and physical properties. Trends in atomic radius The atomic radius increases down the group. Notice that beryllium has a particularly small radius compared with the rest of the group. The radius of an atom is governed by the number of layers of electrons around the nucleus the attraction the outer electrons feel from the nucleus. Compare the electron configurations of beryllium and magnesium: The positive charge on the nucleus is balanced by the negativeness of the inner electrons. This is true for all the atoms in Group 2. More layers of electrons take up more space, due to electron-electron repulsion, so atoms are larger down the group. Trends in first ionization energy The first ionization energy is the energy needed to remove the most loosely held electron from each of one mole of gaseous atoms to make one mole of singly charged gaseous ions. In other words, for 1 mole of reactant in the process shown below: Ionization energy is governed by three factors: The increase in nuclear charge down the group is exactly offset by the increase in the number of inner electrons. However, down the group, the distance between the nucleus and the outer electrons increases and so they become easier to remove—the ionization energy decreases. Electronegativity Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons. It is usually measured on the Pauling scale, on which the most electronegative element fluorine is assigned an electronegativity of 4. Imagine a bond between a magnesium atom and a chlorine atom. At this point, the electrons are so close to the chlorine that there is essentially a transfer of an electron to the chlorine; in other words, ions are formed. In this case, the electron pair is not close enough to the chlorine for an ionic bond to be formed. Because of its small size, beryllium forms covalent bonds, not ionic bonds. The attraction between the beryllium nucleus and a bonding pair is always too great for ionic bonding. In other words, down the group, the elements become less electronegative. Moving down the group, the bonds formed between these elements and other elements such as chlorine become more ionic. Boiling points cyan bars: There is no meaningful trend in boiling points for the Group 2 elements. Atomization energy orange bars: As with boiling points, there is no simple pattern for atomization energies for the Group 2 elements. The chart looks similar, but not identical to, the boiling point chart. In this section, the most common explanation is given and refuted. The Incorrect Explanation Group 2 elements are held together by metallic bonds. The melting points decrease down the group because the metallic bonds are weaker for the larger elements. The deviation of magnesium requires an additional explanation. The atoms in a metal are held together by the attraction between the nuclei and the delocalized electrons. As the atoms increase in size, the nuclei are further away from these delocalized electrons; therefore, the attractions decrease. The atoms are more easily separated to form a liquid, and then a gas. Moving down the group, the atoms form different crystal structures in the solid phase. Beryllium and magnesium each have a hexagonal close-packed structure; calcium and strontium form face-centered cubic structures; the structure of barium is body-centered cubic. There is a change in crystal structure between magnesium and calcium, and it could be supposed that this change accounts for the behavior of magnesium. A problem with this explanation Despite the fact that the first four elements have two different structures, those structures are both coordinated. Each atom is adjacent to 12 atoms. Therefore, the metallic bond should be similar in each case, because the orbitals are going to overlap and delocalize in the same manner.. Any differences due exclusively to structural differences should be minor. By contrast, barium is 8-coordinated like the Group 1 metals. This is a less efficient packing system, which is intuitively expected to result in a much weaker metallic bond. This is obviously not the case; therefore, the change in structure between magnesium and calcium does not explain the deviation. The strength of the metallic bonds Melting points are not accurate indicators of the strength of the metallic bonds. When a metal melts, the bonds are not completely broken—only loosened enough for the

atoms to move around. Metallic bonds are still present in the molten metal, breaking only when the metal boils. Both boiling point and atomization energy are much better indicators of metallic bond strength. Cotton and Wilkinson, in their classic degree level book Advanced Inorganic Chemistry, support this conclusion: Neither is the explanation about weaker metallic bonds down the group accurate. The figures for Group 1 show that the trends for all the properties discussed melting point, boiling point and atomization energy are consistent down the group. Something unique to Group 2 causes complications in simple trends. A final comment There is no convincing explanation for the low melting point of magnesium, or the lack of any pattern with the other two properties. Sharpe, in his degree level book Inorganic Chemistry admits that there is no easy explanation for the variations in the physical data in Group 2.