

= Alkali metal =

The alkali metals are a group (column) in the periodic table consisting of the chemical elements lithium (Li), sodium (Na), potassium (K), rubidium (Rb), caesium (Cs), and francium (Fr). This group lies in the s-block of the periodic table of elements as all alkali metals have their outermost electron in an s orbital : this element / electron configuration results in their characteristic properties . The alkali metals provide the best example of group trends in properties in the periodic table , with elements exhibiting well characterized homologous behaviour .

The alkali metals have very similar properties : they are all shiny , soft , highly reactive metals at standard temperature and pressure and readily lose their outermost electron to form cations with charge + 1 . They can all be cut easily with a knife due to their softness , exposing a shiny surface that tarnishes rapidly in air due to oxidation by atmospheric moisture and oxygen . Because of their high reactivity , they must be stored under oil to prevent reaction with air , and are found naturally only in salts and never as the free element . In the modern IUPAC nomenclature , the alkali metals comprise the group 1 elements , excluding hydrogen (H), which is nominally a group 1 element but not normally considered to be an alkali metal as it rarely exhibits behaviour comparable to that of the alkali metals . All the alkali metals react with water , with the heavier alkali metals reacting more vigorously than the lighter ones .

All of the discovered alkali metals occur in nature : in order of abundance , sodium is the most abundant , followed by potassium , lithium , rubidium , caesium , and finally francium , which is very rare due to its extremely high radioactivity ; francium occurs only in traces , the product of natural decay chains . Experiments have been conducted to attempt the synthesis of ununennium (Uue) , which is likely to be the next member of the group , but they have all met with failure . However , ununennium may not be an alkali metal due to relativistic effects , which are predicted to have a large influence on the chemical properties of superheavy elements ; even if it does turn out to be an alkali metal , it is predicted to have some differences in physical and chemical properties from its lighter homologues .

Most alkali metals have many different applications . One of the best known applications of the pure elements the use of rubidium and caesium in atomic clocks , of which caesium atomic clocks are the most accurate and precise representation of time . A common application of the compounds of sodium is the sodium vapour lamp , which emits very efficient light . Table salt , or sodium chloride , has been used since antiquity . Sodium and potassium are also essential elements , having major biological roles as electrolytes , and although the other alkali metals are not essential , they also have various effects on the body , both beneficial and harmful .

= = Properties = =

= = = Physical and chemical = = =

The physical and chemical properties of the alkali metals can be readily explained by their having an ns¹ valence electron configuration , which results in weak metallic bonding . Hence , all the alkali metals are soft and have low densities , melting and boiling points , as well as heats of sublimation , vaporisation , and dissociation They all crystallize in the body centered cubic crystal structure , and have distinctive flame colours because their outer s electron is very easily excited . The ns¹ configuration also results in the alkali metals having very large atomic and ionic radii , as well as high thermal and electrical conductivity . Their chemistry is dominated by the loss of their lone valence electron to form the + 1 oxidation state , due to the ease of ionizing this electron and the very high second ionization energy . Most of the chemistry has been observed only for the first five members of the group . The chemistry of francium is not well established due to its extreme radioactivity ; thus , the presentation of its properties here is limited .

The alkali metals are more similar to each other than the elements in any other group are to each other . For instance , when moving down the table , all known alkali metals show increasing atomic

radius, decreasing electronegativity, increasing reactivity, and decreasing melting and boiling points as well as heats of fusion and vaporisation. In general, their densities increase when moving down the table, with the exception that potassium is less dense than sodium. One of the very few properties of the alkali metals that does not display a very smooth trend is their reduction potentials: lithium's value is anomalous, being more negative than the others. This is because the Li^+ ion has a very high hydration energy in the gas phase: though the lithium ion disrupts the structure of water significantly, causing a higher change in entropy, this high hydration energy is enough to make the reduction potentials indicate it as being the most electropositive alkali metal, despite the difficulty of ionizing it in the gas phase.

The stable alkali metals are all silver-coloured metals except for caesium, which has a golden tint: it is one of only three metals that are clearly coloured (the other two being copper and gold). Additionally, the heavy alkaline earth metals calcium, strontium, and barium, as well as the divalent lanthanides europium and ytterbium, are pale yellow, though the colour is much less prominent than it is for caesium. Their lustre tarnishes rapidly in air due to oxidation. They all crystallize in the body-centred cubic crystal structure, and have distinctive flame colours because their outer s electron is very easily excited.

All the alkali metals are highly reactive and are never found in elemental forms in nature. Because of this, they are usually stored in mineral oil or kerosene (paraffin oil). They react aggressively with the halogens to form the alkali metal halides, which are white ionic crystalline compounds that are all soluble in water except lithium fluoride (LiF). The alkali metals also react with water to form strongly alkaline hydroxides and thus should be handled with great care. The heavier alkali metals react more vigorously than the lighter ones; for example, when dropped into water, caesium produces a larger explosion than potassium. The alkali metals have the lowest first ionisation energies in their respective periods of the periodic table because of their low effective nuclear charge and the ability to attain a noble gas configuration by losing just one electron. The second ionisation energy of all of the alkali metals is very high as it is in a full shell that is also closer to the nucleus; thus, they almost always lose a single electron, forming cations. The alkalides are an exception: they are unstable compounds which contain alkali metals in a -1 oxidation state, which is very unusual as before the discovery of the alkalides, the alkali metals were not expected to be able to form anions and were thought to be able to appear in salts only as cations. The alkalide anions have filled s subshells, which gives them more stability and allows them to exist. All the stable alkali metals except lithium are known to be able to form alkalides, and the alkalides have much theoretical interest due to their unusual stoichiometry and low ionisation potentials. Alkalides are chemically similar to the electrides, which are salts with trapped electrons acting as anions. A particularly striking example of an alkalide is "inverse sodium hydride", $\text{H}^- + \text{Na}^+$ (both ions being complexed), as opposed to the usual sodium hydride, $\text{Na} + \text{H}$: it is unstable in isolation, due to its high energy resulting from the displacement of two electrons from hydrogen to sodium, although several derivatives are predicted to be metastable or stable.

In aqueous solution, the alkali metal ions form aqua ions of the formula $[\text{M}(\text{H}_2\text{O})_n]^+$, where n is the solvation number. Their coordination numbers and shapes agree well with those expected from their ionic radii. In aqueous solution the water molecules directly attached to the metal ion are said to belong to the first coordination sphere, also known as the first, or primary, solvation shell. The bond between a water molecule and the metal ion is a dative covalent bond, with the oxygen atom donating both electrons to the bond. Each coordinated water molecule may be attached by hydrogen bonds to other water molecules. The latter are said to reside in the second coordination sphere. However, for the alkali metal cations, the second coordination sphere is not well defined as the $+1$ charge on the cation is not high enough to polarize the water molecules in the primary solvation shell enough for them to form strong hydrogen bonds with those in the second coordination sphere, producing a more stable entity. The solvation number for Li^+ has been experimentally determined to be 4, forming the tetrahedral $[\text{Li}(\text{H}_2\text{O})_4]^+$: while solvation numbers of 3 to 6 have been found for lithium aqua ions, solvation numbers less than 4 may be the result of the formation of contact ion pairs, and the higher solvation numbers may be interpreted in terms of water molecules that approach $[\text{Li}(\text{H}_2\text{O})_4]^+$ through a face of the

tetrahedron , though molecular dynamic simulations may indicate the existence of an octahedral hexaaqua ion . There are also probably six water molecules in the primary solvation sphere of the sodium ion , forming the octahedral $[\text{Na}(\text{H}_2\text{O})_6]^+$ ion . While it was previously thought that the heavier alkali metals also formed octahedral hexaaqua ions , it has since been found that potassium and rubidium probably form the $[\text{K}(\text{H}_2\text{O})_8]^+$ and $[\text{Rb}(\text{H}_2\text{O})_8]^+$ ions , which have the square antiprismatic structure , and that caesium forms the 12 coordinate $[\text{Cs}(\text{H}_2\text{O})_{12}]^+$ ion .

===== Lithium =====

The chemistry of lithium shows several differences from that of the rest of the group as the small Li^+ cation polarises anions and gives its compounds a more covalent character . Lithium and magnesium have a diagonal relationship due to their similar atomic radii , so that they show some similarities . For example , lithium forms a stable nitride , a property common among all the alkaline earth metals (magnesium 's group) but unique among the alkali metals . In addition , among their respective groups , only lithium and magnesium form covalent organometallic compounds (e.g. LiMe and MgMe_2) .

Lithium fluoride is the only alkali metal halide that is not soluble in water , and lithium hydroxide is the only alkali metal hydroxide that is not deliquescent . Conversely , lithium perchlorate and other lithium salts with large anions that cannot be polarized are much more stable than the analogous compounds of the other alkali metals , probably because Li^+ has a high solvation energy . This effect also means that most simple lithium salts are commonly encountered in hydrated form , because the anhydrous forms are extremely hygroscopic : this allows salts like lithium chloride and lithium bromide to be used in dehumidifiers and air conditioning .

===== Francium =====

Francium is also predicted to show some differences due to its high atomic weight , causing its electrons to travel at considerable fractions of the speed of light and thus making relativistic effects more prominent . In contrast to the trend of decreasing electronegativities and ionisation energies of the alkali metals , francium 's electronegativity and ionisation energy are predicted to be higher than caesium 's due to the relativistic stabilisation of the 7s electrons ; also , its atomic radius is expected to be abnormally low . All known physical properties of francium also deviate from the clear trends going from lithium to caesium , such as the first ionisation energy , electron affinity , and anion polarizability . The CsFr molecule is also polarized as $\text{Cs}^+ \text{Fr}^-$, showing that the 7s subshell of francium is much more strongly affected by relativistic effects than the 6s subshell of caesium . Additionally , francium superoxide (FrO_2) is expected to have covalent character , unlike the other alkali metal superoxides , because of bonding contributions from the 6p electrons of francium .

===== Nuclear =====

All the alkali metals have odd atomic numbers ; hence , their isotopes must be either odd-odd (both proton and neutron number are odd) or odd-even (proton number is odd , but neutron number is even) . Odd-odd nuclei have even mass numbers , whereas odd-even nuclei have odd mass numbers . Odd-odd primordial nuclides are rare because most odd-odd nuclei are highly unstable with respect to beta decay , because the decay products are even-even , and are therefore more strongly bound , due to nuclear pairing effects .

Due to the great rarity of odd-odd nuclei , almost all the primordial isotopes of the alkali metals are odd-even (the exceptions being the light stable isotope lithium-6 and the long lived radioisotope potassium-40) . For a given odd mass number , there can be only a single beta stable nuclide , since there is not a difference in binding energy between even-odd and odd-even comparable to that between even-even and odd-odd , leaving other nuclides of the same mass number (isobars) free to beta decay toward the lowest odd-even mass nuclide . An effect of the instability of an odd number of either type of nucleons is that odd-numbered elements , such

as the alkali metals, tend to have fewer stable isotopes than even Z - N numbered elements. Of the 26 monoisotopic elements that have only a single stable isotope, all but one have an odd atomic number and all but one also have an even number of neutrons. Beryllium is the single exception to both rules, due to its low atomic number.

All of the alkali metals except lithium and caesium have at least one naturally occurring radioisotope: sodium $Z=22$ and sodium $Z=24$ are trace radioisotopes produced cosmogenically, potassium $Z=40$ and rubidium $Z=87$ have very long half-lives and thus occur naturally, and all isotopes of francium are radioactive. Caesium was also thought to be radioactive in the early 20th century, although it has no naturally occurring radioisotopes. (Francium had not been discovered yet at that time.) The natural radioisotope of potassium, potassium $Z=40$, makes up about 0.012% of natural potassium, and thus natural potassium is weakly radioactive. This natural radioactivity became a basis for a mistaken claim of the discovery for element 87 (the next alkali metal after caesium) in 1925.

Caesium $Z=137$, with a half-life of 30.17 years, is one of the two principal medium-lived fission products, along with strontium $Z=90$, which are responsible for most of the radioactivity of spent nuclear fuel after several years of cooling, up to several hundred years after use. It constitutes most of the radioactivity still left from the Chernobyl accident. ^{137}Cs undergoes high-energy beta decay and eventually becomes stable barium $Z=137$. It is a strong emitter of gamma radiation. ^{137}Cs has a very low rate of neutron capture and cannot be feasibly disposed of in this way, but must be allowed to decay. ^{137}Cs has been used as a tracer in hydrologic studies, analogous to the use of tritium. Small amounts of caesium $Z=134$ and caesium $Z=137$ were released into the environment during nearly all nuclear weapon tests and some nuclear accidents, most notably the Goiânia accident and the Chernobyl disaster. As of 2005, caesium $Z=137$ is the principal source of radiation in the zone of alienation around the Chernobyl nuclear power plant.

== Periodic trends ==

The alkali metals are more similar to each other than the elements in any other group are to each other. For instance, when moving down the table, all known alkali metals show increasing atomic radius, decreasing electronegativity, increasing reactivity, and decreasing melting and boiling points as well as heats of fusion and vaporisation. In general, their densities increase when moving down the table, with the exception that potassium is less dense than sodium.

== Atomic and ionic radii ==

The atomic radii of the alkali metals increase going down the group. Because of the shielding effect, when an atom has more than one electron shell, each electron feels electric repulsion from the other electrons as well as electric attraction from the nucleus. In the alkali metals, the outermost electron only feels a net charge of $+1$, as some of the nuclear charge (which is equal to the atomic number) is cancelled by the inner electrons; the number of inner electrons of an alkali metal is always one less than the nuclear charge. Therefore, the only factor which affects the atomic radius of the alkali metals is the number of electron shells. Since this number increases down the group, the atomic radius must also increase down the group.

The ionic radii of the alkali metals are much smaller than their atomic radii. This is because the outermost electron of the alkali metals is in a different electron shell than the inner electrons, and thus when it is removed the resulting atom has one fewer electron shell and is smaller. Additionally, the effective nuclear charge has increased, and thus the electrons are attracted more strongly towards the nucleus and the ionic radius decreases.

== First ionisation energy ==

The first ionisation energy of an element or molecule is the energy required to move the most

loosely held electron from one mole of gaseous atoms of the element or molecules to form one mole of gaseous ions with electric charge + 1 . The factors affecting the first ionisation energy are the nuclear charge , the amount of shielding by the inner electrons and the distance from the most loosely held electron from the nucleus , which is always an outer electron in main group elements . The first two factors change the effective nuclear charge the most loosely held electron feels . Since the outermost electron of alkali metals always feels the same effective nuclear charge (+ 1) , the only factor which affects the first ionisation energy is the distance from the outermost electron to the nucleus . Since this distance increases down the group , the outermost electron feels less attraction from the nucleus and thus the first ionisation energy decreases . (This trend is broken in francium due to the relativistic stabilization and contraction of the 7s orbital , bringing francium 's valence electron closer to the nucleus than would be expected from non @-@ relativistic calculations . This makes francium 's outermost electron feel more attraction from the nucleus , increasing its first ionisation energy slightly beyond that of caesium .)

The second ionisation energy of the alkali metals is much higher than the first as the second @-@ most loosely held electron is part of a fully filled electron shell and is thus difficult to remove .

== = Reactivity == =

The reactivities of the alkali metals increase going down the group . This is the result of a combination of two factors : the first ionisation energies and atomisation energies of the alkali metals . Because the first ionisation energy of the alkali metals decreases down the group , it is easier for the outermost electron to be removed from the atom and participate in chemical reactions , thus increasing reactivity down the group . The atomisation energy measures the strength of the metallic bond of an element , which falls down the group as the atoms increase in radius and thus the metallic bond must increase in length , making the delocalised electrons further away from the attraction of the nuclei of the heavier alkali metals . Adding the atomisation and first ionisation energies gives a quantity closely related to (but not equal to) the activation energy of the reaction of an alkali metal with another substance . This quantity decreases going down the group , and so does the activation energy ; thus , chemical reactions can occur faster and the reactivity increases down the group .

== = Electronegativity == =

Electronegativity is a chemical property that describes the tendency of an atom or a functional group to attract electrons (or electron density) towards itself . If the bond between sodium and chlorine in sodium chloride were covalent , the pair of shared electrons would be attracted to the chlorine because the effective nuclear charge on the outer electrons is + 7 in chlorine but is only + 1 in sodium . The electron pair is attracted so close to the chlorine atom that they are practically transferred to the chlorine atom (an ionic bond) . However , if the sodium atom was replaced by a lithium atom , the electrons will not be attracted as close to the chlorine atom as before because the lithium atom is smaller , making the electron pair more strongly attracted to the closer effective nuclear charge from lithium . Hence , the larger alkali metal atoms (further down the group) will be less electronegative as the bonding pair is less strongly attracted towards them .

Because of the higher electronegativity of lithium , some of its compounds have a more covalent character . For example , lithium iodide (LiI) will dissolve in organic solvents , a property of most covalent compounds . Lithium fluoride (LiF) is the only alkali halide that is not soluble in water , and lithium hydroxide (LiOH) is the only alkali metal hydroxide that is not deliquescent .

== = Melting and boiling points == =

The melting point of a substance is the point where it changes state from solid to liquid while the boiling point of a substance (in liquid state) is the point where the vapour pressure of the liquid equals the environmental pressure surrounding the liquid and all the liquid changes state to gas . As

a metal is heated to its melting point , the metallic bonds keeping the atoms in place weaken so that the atoms can move around , and the metallic bonds eventually break completely at the metal 's boiling point . Therefore , the falling melting and boiling points of the alkali metals indicate that the strength of the metallic bonds of the alkali metals decreases down the group . This is because metal atoms are held together by the electromagnetic attraction from the positive ions to the delocalised electrons . As the atoms increase in size going down the group (because their atomic radius increases) , the nuclei of the ions move further away from the delocalised electrons and hence the metallic bond becomes weaker so that the metal can more easily melt and boil , thus lowering the melting and boiling points . (The increased nuclear charge is not a relevant factor due to the shielding effect .)

== Density ==

The alkali metals all have the same crystal structure (body @-@ centred cubic) and thus the only relevant factors are the number of atoms that can fit into a certain volume and the mass of one of the atoms , since density is defined as mass per unit volume . The first factor depends on the volume of the atom and thus the atomic radius , which increases going down the group ; thus , the volume of an alkali metal atom increases going down the group . The mass of an alkali metal atom also increases going down the group . Thus , the trend for the densities of the alkali metals depends on their atomic weights and atomic radii ; if figures for these two factors are known , the ratios between the densities of the alkali metals can then be calculated . The resultant trend is that the densities of the alkali metals increase down the table , with an exception at potassium . Due to having the lowest atomic weight of all the elements in their period and having the largest atomic radius for their periods , the alkali metals are the least dense metals in the periodic table . Lithium , sodium , and potassium are the only three metals in the periodic table that are less dense than water : in fact , lithium is the least dense known solid at room temperature .

== Compounds ==

== Hydroxides ==

All the alkali metals react vigorously or explosively with cold water , producing an aqueous solution of a strongly basic alkali metal hydroxide and releasing hydrogen gas . This reaction becomes more vigorous going down the group : lithium reacts steadily with effervescence , but sodium and potassium can ignite and rubidium and caesium sink in water and generate hydrogen gas so rapidly that shock waves form in the water that may shatter glass containers . When an alkali metal is dropped into water , it produces an explosion , of which there are two separate stages . The metal reacts with the water first , breaking the hydrogen bonds in the water and producing hydrogen gas ; this takes place faster for the more reactive heavier alkali metals . Second , the heat generated by the first part of the reaction often ignites the hydrogen gas , causing it to burn explosively into the surrounding air . This secondary hydrogen gas explosion produces the visible flame above the bowl of water , lake or other body of water , not the initial reaction of the metal with water (which tends to happen mostly under water) . The alkali metal hydroxides are the most basic known hydroxides .

Recent research has suggested that the explosive behavior of alkali metals in water is driven by a Coulomb explosion rather than solely by rapid generation of hydrogen itself . All alkali metals melt as a part of the reaction with water . Water molecules ionize the bare metallic surface of the liquid metal , leaving a positively charged metal surface and negatively charged water ions . The attraction between the charged metal and water ions will rapidly increase the surface area , causing an exponential increase of ionization . When the repulsive forces within the liquid metal surface exceeds the forces of the surface tension , it vigorously explodes .

== Compounds with the group 14 elements ==

