Potassium

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Potassium is a chemical element with symbol **K** (derived from Neo-Latin, *kalium*) and atomic number 19. It was first isolated from potash, the ashes of plants, from which its name derives. In the periodic table, potassium is one of the alkali metals. All of the alkali metals have a single valence electron in the outer electron shell, which is easily removed to create an ion with a positive charge – a cation, which combines with anions to form salts. Potassium in nature occurs only in ionic salts. Elemental potassium is a soft silvery-white alkali metal that oxidizes rapidly in air and reacts vigorously with water, generating sufficient heat to ignite hydrogen emitted in the reaction and burning with a lilac-colored flame. It is found dissolved in sea water (which is 0.04% potassium by weight[4][5]), and is part of many minerals.

Naturally occurring potassium is composed of three isotopes, of which ^{40}K is radioactive. Traces of ^{40}K are found in all potassium, and it is the most common radioisotope in the human body.

Potassium is chemically very similar to sodium, the previous element in Group 1 of the periodic table. They have a similar ionization energy, which allows for each atom to give up its sole outer electron. That they are different elements that combine with the same anions to make similar salts was suspected in 1702,[6] and was proven in 1807 using electrolysis.

Most industrial applications of potassium exploit the high solubility in water of potassium compounds, such as potassium soaps. Heavy crop production rapidly depletes the soil of potassium, and this can be remedied with agricultural fertilizers containing potassium, accounting for 95% of global potassium chemical production.[7]

Potassium ions are necessary for the function of all living cells. The transfer of potassium ions through nerve cell membranes is necessary for normal nerve transmission; potassium depletion can result in numerous abnormalities, including an abnormal heart rhythm and various electrocardiographic (ECG) abnormalities. Fresh fruits and vegetables are good dietary sources of potassium. The body

Potassium, $_{19}K$

Potassium pearls under paraffin oil. The large pearl measures 0.5 cm.

Spectral lines of potassium

General properties	
Name, symbol	potassium, K
Appearance	silvery gray
Potassium in the periodic table	
Atomic number (*Z*)	19
Group, block	group 1 (alkali metals), s-block
Period	period 4
Element category	□ alkali metal
Standard atomic weight (±) (A_r)	39.0983(1)[1]
Electron configuration	[Ar] $4s^1$
per shell	2, 8, 8, 1

responds to the influx of dietary potassium, which raises serum potassium levels, with a shift of potassium from outside to inside cells and an increase in potassium excretion by the kidneys.

Properties

Physical

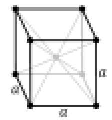

The flame test of potassium.

Potassium is the second least dense metal after lithium. It is a soft solid with a low melting point, and can be easily cut with a knife. Freshly cut potassium is silvery in appearance, but it begins to tarnish toward gray immediately on exposure to air.[8] In a flame test, potassium and its compounds emit a lilac color with a peak emission wavelength of 766.5 nanometers.[9]

Chemical

Neutral potassium atoms have 19 electrons, one more than the extremely stable configuration of the noble gas argon. Because of this and its low first ionization energy of 418.8 kJ/mol, the potassium atom is much more likely to lose the last electron and acquire a positive charge than to gain one and acquire a negative charge (though negatively charged alkalide K^{-} ions are not impossible).[10][11] This process requires so little energy that potassium is readily oxidized by atmospheric oxygen. In contrast, the second ionization energy is very high (3052 kJ/mol), because removal of two electrons breaks the stable noble gas electronic configuration (the configuration of the inert argon).[11] Potassium therefore does not readily form compounds with the oxidation state of +2 or higher.[10]

Potassium is an extremely active metal that reacts violently with oxygen and water in air. With oxygen it forms potassium peroxide, and with water potassium forms potassium hydroxide. The reaction of

Physical properties	
Phase	solid
Melting point	336.7 K (63.5 °C, 146.3 °F)
Boiling point	1032 K (759 °C, 1398 °F)
Density near r.t.	0.862 g/cm^3
when liquid, at m.p.	0.828 g/cm^3
Critical point	2223 K, 16 MPa[2]
Heat of fusion	2.33 kJ/mol
Heat of vaporization	76.9 kJ/mol
Molar heat capacity	29.6 J/(mol·K)
Atomic properties	
Oxidation states	**+1**, −1 (a strongly basic oxide)
Electronegativity	Pauling scale: 0.82
Ionization energies	1st: 418.8 kJ/mol 2nd: 3052 kJ/mol 3rd: 4420 kJ/mol (more)
Atomic radius	empirical: 227 pm
Covalent radius	203±12 pm
Van der Waals radius	275 pm
Miscellanea	
Crystal structure	body-centered cubic (bcc) 
Speed of sound thin rod	2000 m/s (at 20 °C)

potassium with water is dangerous because of its violent exothermic character and the production of hydrogen gas. Hydrogen reacts again with atmospheric oxygen, producing water, which reacts with the remaining potassium. This reaction requires only traces of water; because of this, potassium and the liquid sodium-potassium — NaK — are potent desiccants that can be used to dry solvents prior to distillation.[12]

Because of the sensitivity of potassium to water and air, reactions with other elements are possible only in an inert atmosphere such as argon gas using air-free techniques. Potassium does not react with most hydrocarbons such as mineral oil or kerosene.[13] It readily dissolves in liquid ammonia, up to 480 g per 1000 g of ammonia at 0 °C. Depending on the concentration, the ammonia solutions are blue to yellow, and their electrical conductivity is similar to that of liquid metals. In a pure solution, potassium slowly reacts with ammonia to form KNH_2, but this reaction is accelerated by minute amounts of transition metal salts.[14] Because it can reduce the salts to the metal, potassium is often used as the reductant in the preparation of finely divided metals from their salts by the Rieke method.[15] For example, the preparation of magnesium by this method employs potassium as the reductant:

$$MgCl_2 + 2\,K \rightarrow Mg + 2\,KCl$$

Compounds

The only common oxidation state for potassium is +1. Potassium metal is a powerful reducing agent that is easily oxidized to the monopositive cation, K^+. Once oxidized, it is very stable and difficult to reduce back to the metal.[10]

Potassium hydroxide reacts readily with carbon dioxide to produce potassium carbonate, and is used to remove traces of the gas from air. In general, potassium compounds have excellent water solubility, owing to the high hydration energy of the K^+ ion. The potassium ion is colorless in water and is very difficult to precipitate; possible precipitation methods include reactions with sodium tetraphenylborate, hexachloroplatinic acid, and sodium cobaltinitrite.[13]

Thermal expansion	83.3 µm/(m·K) (at 25 °C)
Thermal conductivity	102.5 W/(m·K)
Electrical resistivity	72 nΩ·m (at 20 °C)
Magnetic ordering	paramagnetic[3]
Young's modulus	3.53 GPa
Shear modulus	1.3 GPa
Bulk modulus	3.1 GPa
Mohs hardness	0.4
Brinell hardness	0.363 MPa
CAS Number	7440-09-7
History	
Discovery and first isolation	Humphry Davy (1807)

Most stable isotopes of potassium

iso	NA	half-life	DM	DE (MeV)	DP
^{39}K	93.258%	is stable with 20 neutrons			
^{40}K	0.012%	$1.248(3)\times10^9$ y	β^-	1.311	^{40}Ca
			ε	1.505	^{40}Ar
			β^+	1.505	^{40}Ar
^{41}K	6.730%	is stable with 22 neutrons			

Potassium oxidizes faster than most metals and forms oxides with oxygen-oxygen bonds, as do all alkali metals except lithium. Three species are formed during the reaction: potassium oxide, potassium peroxide, and potassium superoxide[16] formed of three different oxygen-based ions: oxide (O^{2-}), peroxide (O_2^{2-}), and superoxide (O_2^-). The last two species, especially the superoxide, are rare and are formed only in reaction with very electropositive metals; these species contain oxygen-oxygen bonds.[14] All potassium-oxygen binary compounds are known to react with water violently, forming potassium hydroxide. This compound is a very strong alkali, and 1.21 kg of it can dissolve in as little as a liter of water.[17][18]

Potassium compounds are typically highly ionic and thus most of them are soluble in water. The main species in water solution are the aquated complexes $[K(H_2O)_n]^+$ where n = 6 and 7.[19] Some of the few poorly soluble potassium salts include potassium tetraphenylborate, potassium hexachloroplatinate, and potassium cobaltinitrite.[13]

Isotopes

There are 24 known isotopes of potassium, three of which occur naturally: ^{39}K (93.3%), ^{40}K (0.0117%), and ^{41}K (6.7%). Naturally occurring ^{40}K has a half-life of 1.250×10^9 years. It decays to stable ^{40}Ar by electron capture or positron emission (11.2%) or to stable ^{40}Ca by beta decay (88.8%).[20] The decay of ^{40}K to ^{40}Ar is the basis of a common method for dating rocks. The conventional K-Ar dating method depends on the assumption that the rocks contained no argon at the time of formation and that all the subsequent radiogenic argon (^{40}Ar) was quantitatively retained. Minerals are dated by measurement of the concentration of potassium and the amount of radiogenic ^{40}Ar that has accumulated. The minerals best suited for dating include biotite, muscovite, metamorphic hornblende, and volcanic feldspar; whole rock samples from volcanic flows and shallow instrusives can also be dated if they are unaltered.[20][21] Apart from dating, potassium isotopes have been used as tracers in studies of weathering and for nutrient cycling studies because potassium is a macronutrient required for life.

^{40}K occurs in natural potassium (and thus in some commercial salt substitutes) in sufficient quantity that large bags of those substitutes can be used as a radioactive source for classroom demonstrations. ^{40}K is the radioisotope with the largest abundance in the body. In healthy animals and people, ^{40}K represents the largest source of radioactivity, greater even than ^{14}C. In a human body of 70 kg mass, about 4,400 nuclei of ^{40}K decay per second.[23] The activity of natural potassium is 31 Bq/g.[24]

Source

- Wikipedia: Potassium (https://en.wikipedia.org/wiki/Potassium)