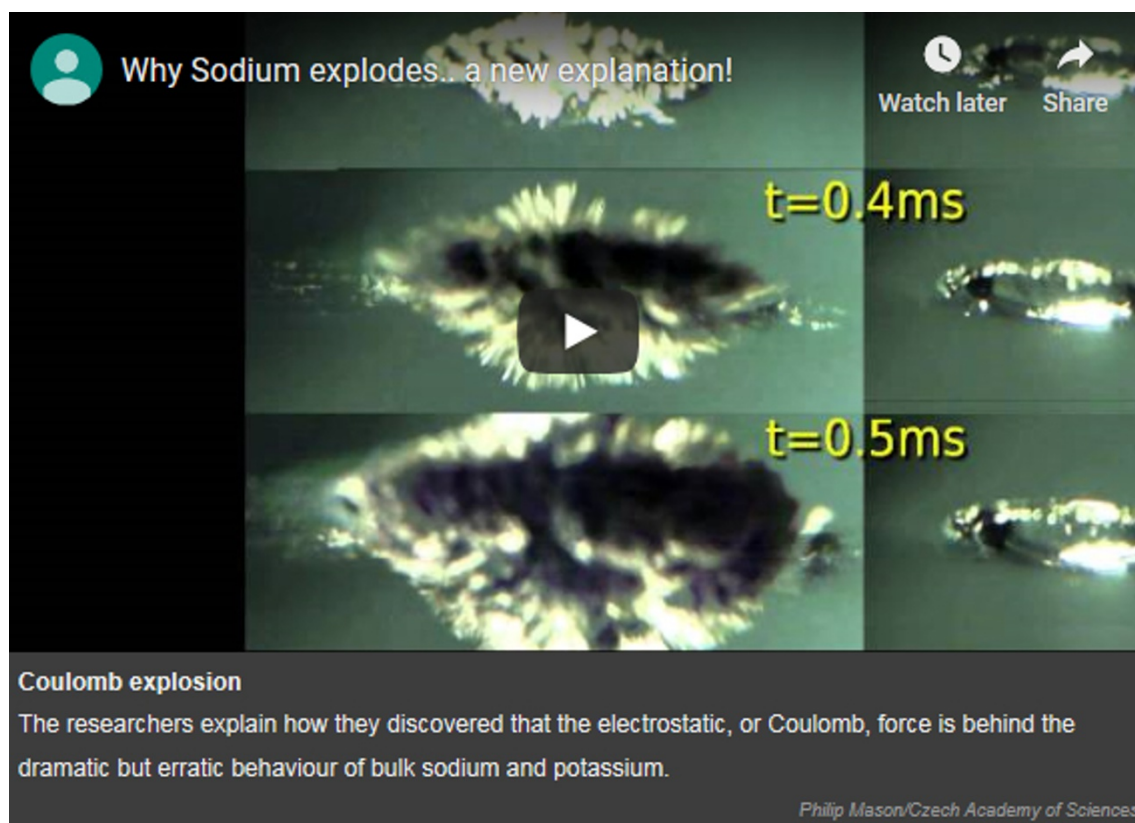


Sodium's explosive secrets revealed

The spectacular reaction of alkali metals with water was poorly understood — despite being a staple of chemistry classes.

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It's the classic piece of chemical tomfoolery: take a lump of sodium or potassium metal, toss it into water and watch the explosion. Although this piece of pyrotechnics has amazed generations of chemistry students, until now it has been misunderstood, a paper in *Nature Chemistry* reveals¹.

The explosion, say Pavel Jungwirth and his collaborators at the Czech Academy of Sciences in Prague, is not merely a consequence of the ignition of the hydrogen gas that the alkali metals release from water. That may happen eventually, but it begins as something far stranger: a rapid exodus of electrons followed by explosion of the metal driven by electrical repulsion.

The neurologist and celebrated author Oliver Sacks, who is also a chemistry enthusiast, described the phenomenon as “chemistry with a vengeance” in his book *Uncle Tungsten*². He wrote of how, as a boy, he and his friends bought a lump of sodium from the local chemicals supplier and dumped it into the Highgate Ponds in North London: “It took fire instantly and sped around and around on the surface like a demented meteor, with a huge sheet of yellow flame above it.”

In bulk, sodium is a silvery metal, a soft solid at room temperature. If it comes in contact with water, it produces sodium hydroxide and hydrogen. Potassium reacts even more vigorously. These reactions unleash a lot of heat, so it seemed straightforward to think that the explosions resulted from the hydrogen igniting.

But for a chemical reaction to have a runaway explosive outcome, the reactants need to mix quickly and efficiently, Jungwirth points out. In the case of the alkali metals, however, the hydrogen gas and steam released at the surface of the metal should prevent more

y of water to fuel it, the reaction should peter out.



Jungwirth's colleague Philip Mason was determined to find out what was really going on — despite the hazards. In one run, he removed his face shield to blow out a small fire, only to have a piece of metal go off in his face, scratching his cheek. To get a reliable reaction — sodium sometimes oxidizes on the surface and will not explode — Mason used an alloy of sodium and potassium that is liquid at room temperature.

High-speed cameras revealed a vital clue to what was fuelling the violent reaction in the early stages. The reaction starts less than a millisecond after the metal droplet, released from a syringe, enters the water. After just 0.4 ms, 'spikes' of metal shoot out from the droplet, too fast to be expelled by heating. What's more, between 0.3 and 0.5 ms, this spiking droplet becomes surrounded by a dark blue/purple colour in the solution.

The reason for these two observations became clear when Jungwirth's colleague Frank Uhlig carried out quantum-mechanical computer simulations of the process with clusters of just 19 sodium atoms. He found that each of the atoms at the surface of the cluster loses an electron within just several picoseconds (10^{-12} s), and that these electrons shoot into the surrounding water, where they are solvated

(surrounded by water molecules)³.

Solvated electrons in water are known to have the deep blue colour briefly observed in the videos. Their departure leaves the metal cluster full of positively charged ions, which repel each other. The result is a so-called Coulomb explosion, in which the cluster bursts apart owing to its ions' mutual electrostatic (Coulombic) repulsion.

"The paper gives a complete and interesting account of the early stages of the reaction", says inorganic chemist James Dye of Michigan State University in East Lansing, an expert in solvated electrons. "I have done the demonstration dozens of times and wondered why sodium globules often danced on the surface, while potassium leads to explosive behaviour."

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References

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3. Young, R. M. & Neumark, D. M. *Chem. Rev.* **112**, 5553–5577 (2012).