

The reaction between sodium and chlorine

Use this experiment to illustrate how the properties of compounds and their constituent elements differ

Kit

- Gas jar containing chlorine (toxic) with lid
- Sodium (highly flammable, corrosive)
- Brick
- Bunsen burner
- Heat resistant mat
- Forceps
- Scalpel or sharp knife
- White tile
- Tissue paper or paper towel

Preparation

Prepare a gas jar full of chlorine. CLEAPSS members should see recipe 24 for guidance on how to do this. The chlorine does not need to be dried before use, but you should make sure the jar has a distinct green colour before performing the demonstration. This indicates that there is a high enough partial pressure of chlorine.

Also, check the mouth of the gas jar is smaller than the brick on which the reaction will be performed, to ensure no chlorine will seep out. Leave the chlorine jar (lidded end down) and the brick (smooth side up) on a heat resistant mat in a fume cupboard for the demonstration.

In front of the class

Students should wear eye protection and the demonstrator should wear goggles. Perform this demonstration in the fume cupboard with the sash down to minimise chlorine escape and eliminate the need for additional safety screens.

Ensure all materials used to handle the sodium are dry. Remove a sample of sodium from the stock bottle, using a pair of forceps, and place it on a tile. Slice off a 6–8 mm (approximately 0.3 g) cube of sodium with a sharp knife, pointing out to

Safety and disposal notes

- Do not attempt to clean the oil off the sodium using solvent – fires have occurred.
- Always keep a container such as a large, labelled, clean soft drink bottle containing clean, dry sand on hand for extinguishing fires when working with alkali metals.
- Stick to the advised size of sodium cube. Larger lumps may only react on the outer surface, leaving sodium below, which can catch fire or explode during clean-up.
- When the brick has cooled after the reaction, you may wish to moisten the area with isopropyl alcohol (propan-2-ol) before submerging the brick in cold water to remove all remaining traces of alkali metal

your students the softness of the metal and its lustre beneath the fresh cut in the process.

Return the rest of the sodium to the stock bottle and use a piece of paper towel to remove as much oil as possible from the sodium cube.

You will need to heat the sodium with the air hole open on the Bunsen, but if the gas tap is fully open, this can blow the sodium around when it melts and forms a ball. To prevent this, you can reduce the gas pressure at the tap. Use caution here as the gas pressure from the Bunsen burner can blow the sodium around on the brick's surface. First the sodium will melt, then any residual oil will burn off and finally the metal itself will ignite. Once alight, remove the lid of the gas jar containing the chlorine and place over the top of the burning metal (it can be helpful to have an extra pair of hands here).

Once the glass jar is in place, the yellow sodium flame intensifies and the white 'smoke' of sodium chloride is produced.

Alternatives

If you don't have a brick, then you can use an upside-down porcelain crucible lid on a small bed of sand as a substitute – but this may break during the demonstration. Do not attempt to use porcelain with other alkali metals.

It's not advised to perform this demonstration with a deflagrating spoon as brown iron chloride smoke (from a steel spoon) or copper chloride smoke (from a brass spoon) will be produced and spoil the effect.

Teaching goal

In my first year of teaching, I used this demonstration alongside a variation of an activity described by Vanessa Kind in her report *Beyond Appearances: Students' misconceptions about basic chemical ideas to introduce structure and bonding*. Using chocolate buttons for electrons, I laid out the electronic configuration of a sodium atom and a chlorine atom on my desk. The students made observations about the forces between the electrons and the nucleus, and the student who noticed that there was 'an extra electron on the sodium atom and a gap in on chlorine' was rewarded with the rest of the chocolate while I went on to explain that ionic compounds form when metals give their outer shell electrons to non-metals. Then, we witnessed this with the demonstration outlined above.

The above description of what happens when sodium and chlorine react is, of course, not true! The ionic bond is simply the electrostatic attraction between oppositely charged ions, and electron transfer models like the one I used as a new teacher are known to lead to learners constructing incorrect 'simple molecular' models of ionic compounds. Such models lead to all kinds of problems with visualising and explaining reactions, especially those in solution.

Currently, I follow a bonding teaching sequence suggested by Keith Taber in his book *Teaching secondary chemistry*. Metals are introduced first to normalise the existence, and indeed prevalence, of ions before we look for them in solutions of ionic substances using conductivity probes, electrolysis and precipitation reactions. Arriving at the ionic model this way illustrates that most of the sodium chloride that students are ever likely to encounter comes not from electron transfer between atomised elements, but rather from the crystallisation of ions which already existed in solution.

So, when do I use this demonstration? I find it sits neatly alongside other reactions of elements for a younger audience where bonding is less important than stressing the differences in the properties of compounds from those of their constituent elements (also see this Exhibition chemistry article on *Clarifying compounds and mixtures*). After all, I'm sure we'd all agree that crisps with sodium and chlorine is an unappetising proposition.