# TEACHING CHEMISTRY WITH DEMONSTRATIONS AND ACTIVITIES

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Let me begin with a confession: although I'm a chemistry lecturer (and enjoying teaching and research in chemistry very much), my first love was biology. As a child, I was fascinated by phenomena like the germination of seeds and metamorphosis of caterpillars, and have kept pets ranging from snails to cats. These were tangible, everyday things that I could see, feel and touch, and get personally involved with.

My first exposure to chemistry was when my primary four teacher mentioned that "limewater turns cloudy when one blows into it" (as an example of change). I was eagerly waiting to see a demonstration of the phenomenon in class, but it never happened. The fact did not make much sense to me in words; I just had to see it for myself. So I went into my backyard and plucked some lime and squeezed the juice out into a glass, took a straw, and started blowing into the lime juice. Of course, the lime juice did not turn cloudy! The conclusion: limewater is probably NOT lime juice. But since my teacher was not prepared to discuss my query, the mystery of limewater was not solved until I was in secondary two. However, as I was personally involved in the process of finding out about limewater, the knowledge that limewater is a saturated solution of $Ca(OH)_2$ in water remain with me.

I was blessed with a secondary school chemistry teacher who was very keen on conducting live demonstrations and hands-on activities in class. I remember the great joy and wonder I had when he burned some sulfur in a jar of oxygen, prepared before our eyes by the $MnO_2$-catalysed decomposition of hydrogen peroxide and collected by downward displacement of water. Hmmm, sulfur burns with a blue flame! Water was then added to the jar, which seemed to be still empty, followed by a piece of blue litmus paper. The litmus paper turned red! Conclusion: the burning of sulfur produces a colourless gas that dissolves in water to give an acidic solution.

That was when I fell in love with chemistry. Henceforth, it didn't matter if I had the worst chemistry teacher in the world: my liking for the subject kept me going like an exothermic reaction generating enough heat to sustain itself.

The point I'm trying to make is that it is very important to make chemistry (and science in general) real and less abstract to students, through demonstrations and hands-on activities. Most students cannot realistically be expected to be motivated by the theoretical knowledge of chemistry alone, and be able to persist with a purely intellectual interest in chemistry (Fensham, 1984). Students also understand and remember things much better by seeing, smelling, and doing, especially when the experiments involved are interesting and spectacular. Chemistry is relatively boring to read and "do problems" about, especially when one has to balance, for example, apparently meaningless equations like "$Cu + 4\,HNO_3 \longrightarrow Cu(NO_3)_2 + 2\,NO_2 + 2\,H_2O$" before being told what are "$HNO_3$", "$Cu(NO_3)_2$", or "$NO_2$". What a difference it would make if students are allowed to watch some copper turnings being consumed by concentrated nitric acid, with dense brown fumes being evolved and the solution turning deeper and deeper blue in the process! The students can also feel the beaker and realise that much heat is given off as well. Even in the IT age, some things are still better appreciated "life" and hands-on!

With that in mind, let me share with you a few of my favourite chemical demonstrations:

## 1. AN ENDOTHERMIC REACTION: A COLD PACK

This is a reaction between two solids, barium hydroxide octahydrate and ammonium chloride:

$$Ba(OH)_2 \bullet 8\,H_2O + 2\,NH_4Cl \longrightarrow BaCl_2 + 2\,NH_3 + 10\,H_2O$$

Weigh out 6.3 g of $Ba(OH)_2 \bullet 8\ H_2O$ and 2.1 g of $NH_4Cl$ in separate beakers. Seal the mouths of the beakers with cling-film if you plan to perform the demonstration more than one hour later. Otherwise, it is sufficient to cover the beakers with watch glasses. Close the bottle of $Ba(OH)_2 \bullet 8\ H_2O$ very tightly after use since the compound reacts slowly with $CO_2$ in the air to form $BaCO_3$.

During the demonstration, pour the contents of both beakers into a "zip-loc" bag (9 × 13 cm) and seal the bag in the usual way after pressing out the air. Rub the bag vigorously between your palms, or ask a student to do it. As the solid mixture in the bag is rubbed, a reaction occurs, and the mixture turns into a liquid due to the formation of water. The mixture will also turn very cold since the reaction is very endothermic. The ammonia gas formed should not escape unless the seal of the "zip-loc" bag is poor (some solid may be trapped in the sealing strip), or if the bag is rubbed so hard that it opens up. Pass the bag around for students to feel.

This activity can be used to introduce the topic of energy changes in reactions. It also shows that reactions can occur between solids.

## 2. FIREWORKS: VERY EXOTHERMIC REDOX REACTIONS

$$C_{12}H_{22}O_{11} + 8\ KClO_3 \longrightarrow 12\ CO_2 + 11\ H_2O + 8\ KCl$$

In fireworks, a combustible material (cane sugar, sucrose, in this case) reacts vigorously with a strong oxidising agent (*e.g.* potassium chlorate), producing much heat and light energy.

Grind about 2 g of sugar (preferably refined sugar) to a fine powder and place it in a dry beaker. Place about 2 g of $KClO_3$ into another dry beaker and gently crush any lumps using your finger (without any open wounds) or a plastic spatula. NEVER grind or hammer $KClO_3$ or it may decompose violently. Mix the two powders thoroughly by shaking them together in one beaker.

During the demonstration, pour the sugar-chlorate mixture onto a large, dry piece of tile (20 × 20 cm), well clear of any personal items or laboratory equipment and chemicals. Darken the room and set off the fireworks by adding one or two drops of concentrated sulfuric acid, holding the dropper at arm length. The powder with ignite with a flare, giving a lilac flame due to the presence of potassium in $KClO_3$. The sulfuric acid reacts vigorously with $KClO_3$ to give off much heat. The heat starts the oxidation reaction, which is even more exothermic and can thus sustain itself.

Interesting variations to the above procedure include the addition of about 1 g of a strontium, calcium, or barium (or copper) salt to the sugar-chlorate mixture. This will colour the fireworks crimson, red, or green, respectively. The sulfuric acid can also be added by dipping a glass rod in a beaker containing a little of the acid (so that the beaker appears empty from a distance) and then touching the powder with the end of the glass rod dipped in the acid. This creates a "magic wand" effect.

Do not store any leftover sugar-chlorate mixture. Flush it down the sink.

This demonstration can be used to introduce the topic of energy changes in reactions and show the effects/danger of strong oxidising agents. The colouring of the fireworks with various salts also serves to reinforce the knowledge of flame tests in qualitative analysis.

## 3. CATALYTIC OXIDATION OF AMMONIA

$$4\ NH_3 + 5\ O_2 \longrightarrow 4\ NO + 6\ H_2O$$

In this demonstration, a coil of platinum wire is heated in a Bunsen flame until it glows, and immediately held over some concentrated ammonia solution. The hot platinum catalyses the air-oxidation of ammonia gas (escaping from the solution) to form nitric oxide and water vapour. This reaction is very exothermic, and the heat given off is sufficient to maintain the wire at red-hot temperature. At this temperature, the catalytic activity of the platinum is also maintained, so that it glows by itself as long as there is a constant supply of ammonia and oxygen. This is the basis of the Ostwald process for the manufacture of nitric acid.

Before the demonstration, prepare the coil by winding a 15-20 cm length of platinum wire around a glass rod. Attach one end of the coiled wire to the glass rod, leaving the other end dangling about 2 cm above the level of the ammonia solution when the rod rests on the mouth of the container. The best container to use, in my experience, is a glass bottle (about 500 mL) with a medium-sized mouth (4-5 cm in diameter). A beaker is not so effective because its wide mouth allows the ammonia to escape too quickly and hence prevents a sufficiently-high concentration of ammonia gas to be built up around the platinum wire. A bottle with too small a mouth does not allow enough air circulation to supply oxygen at a fast enough rate.

As a control experiment, the heated wire can be placed over distilled water instead of ammonia solution. The glow in the wire will disappear very quickly, showing that the presence of ammonia is indeed necessary for heat to be evolved.

This demonstration can be used to illustrate exothermic reactions and catalysis.

## 4. FLOWERING CHRISTMAS TREE: PREPARATION OF SALTS BY PRECIPITATION

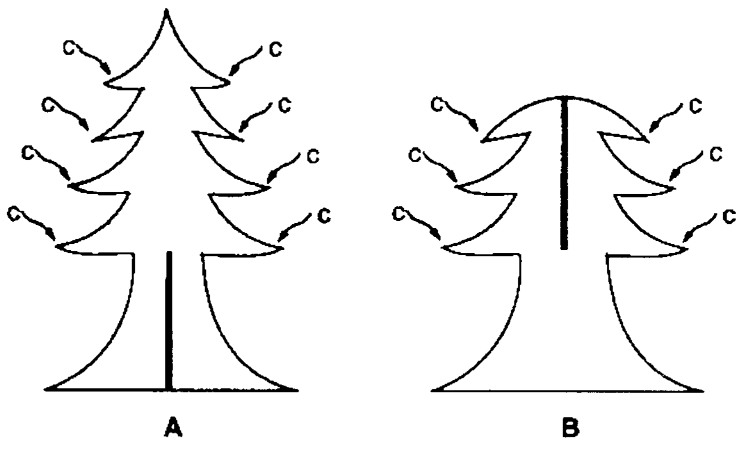
$$2\,M^{+} + SiO_3^{2-} \longrightarrow M_2SiO_3 \downarrow$$

$$M^{2+} + SiO_3^{2-} \longrightarrow MSiO_3 \downarrow$$

$$2\,M^{3+} + 3\,SiO_3^{2-} \longrightarrow M_2(SiO_3)_3 \downarrow$$

We are all quite familiar with the "chemical garden" experiment that illustrates the formation of sparingly soluble salts by precipitation or "double decomposition". Here's a very interesting variation of the same experiment: instead of dropping crystals of different salts into "water glass" (aqueous solution of sodium silicate), immerse pieces of thick art paper, cut into shapes of Christmas trees and impregnated with different coloured salts, in water glass. Crystals of the metal silicates of different colours will "grow" on the "trees"!

The paper used must be very absorbent, so that the water glass can soak through easily. Best effects are obtained using green-coloured art paper cut into the following shapes:




Dip the corners marked "c" into saturated solutions of various salts of different colours (colours must be due to metal cations and not anions) and air dry. Slot **A** onto **B** to form a stable "tree" and immerse the whole tree in water glass in a beaker (stabilise the foot of the tree with plasticine or "blu-tack"). Crystal "flowers" will "bloom" at the corners of the tree after 1 – 2 hours. Very beautiful results are obtained if the set-up is sheltered from vibrations.

## CONCLUDING REMARKS

The above represent just the tip of the iceberg of the demonstrations and activities that can be used in chemistry lessons. Many good books on performing chemical demonstrations are available in the NIE and/or MOE libraries, some of which are given in the list of references below. As a general rule, students should, as far as is practicable, see or do all the important reactions mentioned in the theory lessons. As with any other chemistry experiments, good laboratory safety practice should be observed at all times, and the teacher should always try the demonstration/activity first himself/herself, to make sure it works and to get the feel of it. Effort should also be made to clarify the link between theory and experiment, and to encourage students to carry out their own observations and investigations, to avoid them treating the demonstrations as pure entertainment.

The smooth integration of demonstrations into lessons take some practice to perfect, but the rewards outweigh the initial difficulties. The important thing is to start trying the activities, to get over the inertia. I wish you all the best in your efforts to make chemistry more interesting and meaningful to your students!

## REFERENCES

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