

Chemistry outdoors

Peter Borrows

There is a surprising amount of simple chemistry in the outdoor environment. Including this in your teaching can both enrich it and provide a way of revising familiar concepts in new contexts

Chemistry teachers face a rather daunting challenge. Everybody knows that biology happens outdoors – you only have to look around. Physics, too, is pretty obvious outdoors: think power lines and mobile phones. Chemistry is less obvious. If you define chemistry as a process of change (and this is an oversimplification but the Chinese for chemistry literally translates as ‘change study’), the chemistry is not actually happening at the time you see these artefacts. You only see the product of the chemistry. Whereas if someone is using a mobile phone, microwaves are being emitted and absorbed; the physics is actually happening. It really is important to convince pupils (and some teachers) that chemistry doesn’t just happen in test tubes.

How to use the opportunities

Twenty years ago this journal (Borrows, 1984) introduced the idea of a chemistry trail around Pimlico School in central London. Just like a nature trail, this was a short walk around the local environment, stopping at various places and commenting on or explaining the features of chemical interest. There

was an accompanying booklet. Since then, sometimes with the help of other teachers, I have devised other chemistry trails and others have followed suit. Some of these trails took the form of properly printed leaflets but others were simply a set of brief notes as a reminder to the guide leading a group.

However, a trail has its limitations. It can be time-consuming to organise a group to go out of school and ensure adequate supervision. The trouble is certainly worthwhile if you have somewhere exciting to go, but perhaps not for a trip to the local multi-storey car park. In any case, talking to a group of 30 pupils may be easy in the relative peace of the countryside but can be a nightmare against the roar of high-street traffic. Staying on the school site (if it is chemically rich enough) presents fewer problems but may disrupt classes nearby. So compiling a local chemistry trail can be a useful exercise to raise teachers’ awareness of the potential of the local environment and to build up a valuable resource but, in practice, it may be better to use the components piecemeal.

Pupils naturally walk around the school anyway, and they go down the high street on their way to and from school. Observing some aspect(s) of the chemical environment could be set as a homework assignment, probably a rather more valuable task than many that are set. Of course, there must be due regard to issues surrounding the safety of the pupils. With digital cameras readily available and much better quality printing or photocopying in schools, it is now feasible to take pictures of buildings or artefacts and include them in a worksheet. Pupils can then be told to find the item and answer some questions about it. Alternatively, illustrations might show non-local examples and pupils asked to locate similar examples in their environment. Such tasks can then be infiltrated throughout the curriculum, to help teach, develop and reinforce chemical ideas met at various stages. In a spiral curriculum,

ABSTRACT

Teachers can use the local environment to reinforce important ideas of elementary chemistry. Concrete can be used to discuss acids, bases, indicators and neutralisation, rates of reaction, acid rain, diffusion, thermal decomposition, reversible reactions, composite materials, hazards in everyday life. In churchyards it is usually possible to explore the physical properties of metals, their reactivity with air and water, the rusting and expansion of iron, the action of acids on carbonates, the effect of rate of cooling on the size of crystals, rock types, and so on. These and many other ideas are discussed in this article.

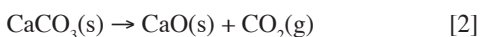
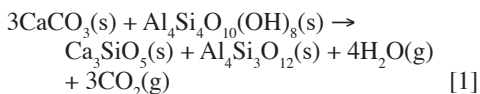
encountering an idea previously met in a different context can be a helpful way of revising ideas and testing understanding before moving on.

Even well-qualified chemistry graduates may find they do not have much factual knowledge of their chemical environment but can readily understand and build on it once the ideas are pointed out. Therefore much of this article takes a few examples and discusses the chemistry of each and suggests ways in which they could be used. The first two of these examples start from the environment (concrete buildings, churchyards) and explore what chemistry can be extracted from them. Others start from a curriculum perspective (e.g. metals, crystals): this is what I am teaching, how can I use the local environment to help me? The two approaches are complementary and some of the examples overlap in parts. In places, italics are used to highlight questions that you can ask the children, but again these are only examples.

What's out there? A concrete example

Consider cement and concrete. Concrete is absolutely everywhere – think flyovers, multi-storey car parks, under-passes, and so on. To the surprise of most people, much of the chemistry of this ubiquitous building material is quite accessible to key stages 3 and 4 (ages 11–16) pupils.

Cement is made by roasting a mixture of crushed calcium carbonate and clay in huge kilns at about 1400 °C. Two main reactions occur:



Reaction [1] is rather beyond key stages 3/4, although gifted and talented pupils may enjoy the challenge of trying to balance the equation. Reaction [2], however, is a good, practical example of thermal decomposition, which appears in most syllabuses. As a result, cement contains some calcium oxide, which is corrosive. Bags of cement nowadays should bear the appropriate safety symbol.



This is a good opportunity to teach pupils about the hazards of everyday life. If they were using something corrosive in school science what safety precautions would they take? Does this happen on a building site? Or in a DIY activity at home?

Concrete is made by mixing cement, sand and small pebbles with water. The pebbles certainly don't react, so concrete is an example of a composite material. Cement, sand and water, without the pebbles, are used to make mortar, for joining bricks and other building materials together. The characteristic of both mortar and concrete is that they set, that is, they gradually harden. There is some complicated chemistry going on here but one of the key reactions is a simple acid/base neutralisation, involving carbon dioxide from the air:



Reaction [3] is the reverse of reaction [2] so there is an opportunity to teach something about reversible reactions. Concrete and mortar are not very porous and hence carbon dioxide will take some time to diffuse from the surface to the interior. Hence the carbonation reaction [3] occurs mainly in the surface layers. As calcium oxide is basic and calcium carbonate is neutral it follows that a lump of concrete is likely to be neutral near the surface and basic below that. This is easily demonstrated with a lump of concrete (you can find lumps lying about all over the place but it is best not done with an actual building!). Pour phenolphthalein solution (HIGHLY FLAMMABLE) on to a lump of concrete. Disappointingly, the colourless solution stays colourless, because the surface layer is neutral. Wearing eye protection, hit your lump of concrete hard with a hammer so that it breaks up. Quickly pour phenolphthalein solution on to a newly exposed surface. Dramatically, a deep pink colour develops, owing to the basic calcium oxide. The colour soon fades as atmospheric carbon dioxide neutralises the base.

How long does the colour take to fade? Could this form an investigation on rates of reaction?

You could, of course, use universal indicator, which would be more familiar to most pupils in key stages 3/4 than phenolphthalein. However, the colour changes are much less dramatic and perhaps it is worth teaching that not all indicators are universal.

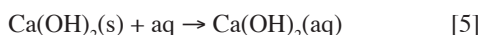
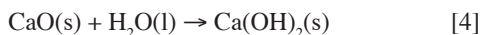
If you look carefully at the pink colour on the newly exposed concrete surface, you may well be able to see there is no pink colour within a few millimetres of the old exposed surface. Carbon dioxide cannot easily diffuse from the surface into the bulk. Hence the carbonation layer is often only a few mm deep. The depth of penetration can be used to date a piece of concrete.



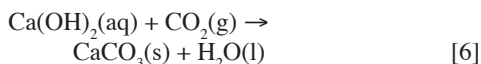
Figure 1 Concrete stalactites formed under a concrete structure.

How might this be used in some forensic science investigations, e.g. to find when a body was entombed?

Water slowly percolates through concrete. Any calcium oxide (quicklime) will react with the water, forming calcium hydroxide (slaked lime, reaction [4]), and this then dissolves to give calcium hydroxide solution (lime water, reaction [5]):



Lime water reacts with carbon dioxide, forming a cloudy suspension of calcium carbonate:



This happens in most concrete structures. Rainwater percolating through the concrete produces calcium hydroxide solution. This may drip down from concrete surfaces, such as arches and bridges. Be warned: this is a fairly concentrated solution of calcium hydroxide, perhaps even saturated, so the pH may well be about 10. Test it with pH paper.

Give the pupils some pH paper, so they can test the drips for homework.

As it drips down, it reacts with atmospheric carbon dioxide leaving a deposit of calcium carbonate. The deposit gets bigger and bigger. A stalactite develops, although by very different reactions to those that form the stalactites in the great show caves. Look up in most concrete structures and you will spot the stalactites growing (see Figure 1).

Can you estimate the rate of growth? One centimetre per year would be typical, as compared with a centimetre every thousand years for cave stalactites.

Have you ever noticed concrete stalactites before? Can you stop looking for them now? Suddenly, car parks become much more interesting. If there are stalactites you may also expect to see stalagmites but as these would usually be where people walk, they get destroyed before they develop much. Just as in caves, you don't get stalactites if the solution cannot form drips. If it seeps out of the walls you may get other formations. (Particularly good, iron-stained, examples can be seen on some of the walls of the escalator hall at Westminster underground station in London.)

Ask the pupils if they can spot concrete stalactites in their environment. Are there any signs of stalagmites or other formations?

There is another consequence of the basic nature of concrete. If it can neutralise carbon dioxide in the air, it can also neutralise other acid gases, for instance acid rain. Lichens are common enough in the countryside but you don't often find them in towns because they don't grow well in acidic conditions. If you do find them, it is likely to be on concrete surfaces. Look out for grey-green or orange patches.

Ask the pupils to find examples of lichens. On what types of surface are they growing?

The chemical potential of churchyards

Churchyards are often a rich source of chemistry. The church building, the gravestones, walls and fences are worth looking at. Most churchyards have been there for a long time and hence even slow chemical reactions have had time to take place to an observable extent.

The metals

You can usually see a range of metals and this can be used to illuminate ideas about the reactivity series as well as the physical properties of metals. Probably the

church tower will have a lightning conductor. Made of copper because of its good electrical conductivity, most are likely to have a green patina, usually said to be basic copper(II) carbonate, $\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2$, although the truth is probably more complicated. You can remind pupils that transition metal compounds are normally coloured. Depending what building material has been used for the church, you may also see some green stains on the stonework. You can make the point that copper cannot be a very reactive metal if you can leave it lying around in the atmosphere but it is obviously not completely inert.

Lead is also very common in churchyards. It finds a number of uses, mainly due to the fact that it is soft metal, easy to work and also unreactive. It is sometimes used for drainpipes, taking the water off church roofs. Greyish in appearance, the surface layer is easily scratched off, with a coin for example, to reveal the shining metal underneath. Lead has also been used sometimes on gravestones. Where the material of the stone is white (e.g. marble), sometimes letters made of lead are placed into the incised letters so that the inscription stands out clearly – almost black on white (see Figure 2).

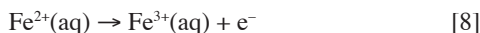
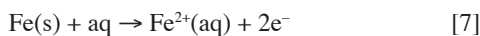


Figure 2 Lead lettering on a gravestone used to make the inscription clearer.

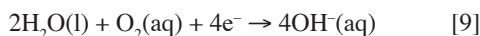


Figure 3 Rusting iron expanding to crack concrete or other stone.

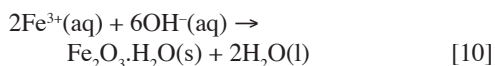
Lead has also been used for another purpose, this time the result of its low melting point (328°C). Quite often, you may find iron fences (around particular graves or around the churchyard itself). The iron posts are often set into holes in some building stone, such as sandstone. The hole is bigger than the iron post and so the post has to be securely fixed into it. In the past, molten lead was used for this purpose. In the long run, this causes problems because you have two metals in contact. There is an electrolytic cell. The more reactive metal (iron, in this case) corrodes. It is often easy to see that the iron gets thinner, especially where it is close to the lead:



The electrons released flow through the iron into the lead. At its surface the electrons are passed to water molecules which in the presence of oxygen form hydroxide ions:



The hydroxide ions diffuse towards the iron(III) ions, where they combine to form hydrated iron(III) oxide, that is, rust:



When iron forms rust, the volume increases substantially. The iron railing, fitted into a hole in the sandstone, no longer fits. The pressure increases and the stone cracks (Figure 3).

Other metals may sometimes be found in churchyards. Look out for galvanised (i.e. zinc-coated) iron. There could well be brass (a zinc/copper alloy) or aluminium. These may be associated with

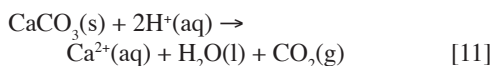
door fittings, boxes for electricity meters or watering cans. Inside the church you may find artefacts of silver or even gold (in Norwich Cathedral, there are rhodium-plated candlesticks).

What metals have been used in a particular churchyard? What physical and chemical properties makes some suitable for these jobs? What evidence is there for the corrosion of these metals? Can the corrosion be slowed down?

The stones

The stones of which the church and the graves are constructed may also tell an interesting story. Granite is often used for gravestones. It is an igneous rock composed of coloured plates of mica, for instance $\text{KAl}_2\text{Si}_3\text{O}_{10}(\text{OH})_2$, small clear crystals of quartz, SiO_2 , and quite large crystals of feldspars, for instance KAlSi_3O_8 . As magma cools, crystals of the three minerals are formed but there is a wide variation in crystal sizes, depending on the rate of cooling. Pupils may learn, using salol, that slower cooling leads to larger crystals. They are thus in a position to explain how some granite rocks have small crystals and others large ones.

Granite, an igneous rock, is unreactive. The lettering carved on the gravestone 150 years ago will be as readable now as it was then. This is much less true of other materials used for gravestones, for example sandstone or limestone, which are both sedimentary rocks. Some sandstones comprise grains of sand (silicon dioxide) held together by a natural cement of calcium carbonate. Calcium carbonate, whether in sandstone or limestone, is susceptible to attack by acid rain:



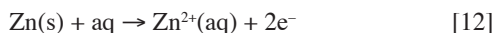
Hence the lettering on sandstone and limestone graves is often unreadable. Gargoyles on the church may similarly lose their detail. Lettering on marble tends to survive better, although it too is calcium carbonate. This is because marble, a metamorphic rock, is much denser and so less susceptible to attack. Since gravestones carry dates, pupils may be able to compare the rate of reaction of the different materials with acid rain. You may well find lichens growing on gravestones but only on those where the material can neutralise acid rain. You will not find them on granite.

What stones have been used for graves in a particular churchyard? Classify the inscriptions from 0 (unreadable) to 5 (sharp and clear). Also record the type of stone and the date of the

grave. What patterns emerge? Which stones show lichen growth?

Teaching some properties of metals

We often want to teach ideas about the reactivity series and we can use the surprising number of metals in most environments to help us. Iron (or steel) can be found everywhere and self-evidently corrodes in the atmosphere. A range of strategies is employed to slow down its corrosion. Copper is appreciably less reactive. See above for copper lightning conductors but also think about statues – or even the pipes in some public toilets. Lead is similar. A dull greying of the surface may be evident but it mostly survives intact. You probably won't find many objects made of gold or silver lying around, except perhaps in churches or jeweller's shop windows, but if you do the gold will be as bright as it ever was and the silver will at most have a surface blackening. On the other hand, zinc is more reactive than iron. That is why iron rubbish bins, railings, and so on, are often coated with a layer of zinc (galvanised). The zinc corrodes in preference to the iron. In fact, there is an electrolytic reaction:



On the surface of the less reactive iron, reaction [9] occurs.

We don't find sodium in the environment: it is far too reactive to survive more than a few minutes unless it is protected from the environment by wrapping it up in glass, as in sodium street lamps. Aluminium is anomalous. It is normally protected from corrosion by an invisibly thin layer of the oxide. However, occasionally you can spot old samples of aluminium in the environment that do show signs of building up a white deposit of the oxide. Chromium, on the other hand, always seems to retain its shine.

If pupils have been taught about the reactivity series, and given a few hints as outlined above, ask them to see what evidence they can find in the environment in support of the series. Alternatively, show them pictures of some local examples and ask them how these pictures support the idea of a reactivity series and what further examples they can find. Ask them to find examples of the different methods used to prevent the corrosion of iron.

Teaching about crystallisation

We teach pupils that the regularity of crystals gives us some evidence about the underlying order on the atomic scale and you can find evidence of crystallisation in most local environments.

We protect iron from rusting by galvanising it (see above). Sheets of iron are dipped into molten zinc. A thin layer of zinc clings to the surface, and this crystallises as it cools. Rapid cooling leads to small crystals; slow cooling gives larger, feathery patterns. Sometimes on the same pieces of galvanised iron there are patches of crystals of very different sizes (see Figure 4).

Other metals also show patterns of crystallisation. When molten metals crystallise, crystallisation starts at many places. The crystals grow into each other, forming interlocking grains. Although these are usually very evident on new samples of zinc, the pattern can become even more evident over time as a result of etching, that is, as acids selectively dissolve atoms at the edge of the grain. I have seen the grains in brass door fittings, etched by generations of sweaty hands, and you may occasionally see grains in aluminium for similar reasons.

Not only metal crystals are easily found in the environment. As magma cools, crystals of the three minerals found in granite are formed; wide variations in crystal sizes can be found, depending on the rate of cooling. Granite is widespread; for example it is used in gravestones, kerbstones and granite setts (cobblestones).

Ask pupils what examples of crystals they can find in the environment. Can they find the same material in different-sized crystals, hence demonstrating different rates of cooling?

Pollution

Children are fascinated to know how polluted their local environment is. You may be able to obtain data from the Internet, for instance on air and river pollution (see websites). Try to visit the local sampling point. You may also be able to relate air pollution to the physical condition of local buildings.

The history of chemistry and chemical industries

In most areas you can find relics of past chemical industries. Place names such as Lime Kiln Lane and Salter's Way have a fairly obvious origin. But place names involving the word Bitterne (e.g. on the Hampshire coast) refer to the ponds where seawater



Figure 4 Zinc crystals.

was evaporated to make salt and Pwlheli (in North Wales) means 'salt pool'. Copperas is a common place name on the coasts of Essex and Kent, and refers, confusingly, to iron sulfide, from which iron sulfate was produced (for tanning and other purposes).

Another historical line to pursue is memorials to famous chemists. These may be statues (as in the statue of Michael Faraday in Leeds), wall plaques (as that to Robert Boyle in Oxford) and graves (as that of Lord Rayleigh, the discoverer of argon, in Essex). There are also rather more quirky ones, such as the plaque in Chester recording the launch of a hydrogen balloon in 1785, which carried the adventurer 25 km.

Ask your pupils to find local examples. How might they have generated so much hydrogen over 200 years ago? Your class should be able to get it right – by dropping iron filings into several barrels of sulfuric acid.

What else is in the chemical environment?

The same environmental observation can sometimes be used to illustrate different chemical concepts or different parts of the science curriculum. For example, a rusting railing can be used to think about the reactivity of iron in comparison with other metals, about methods of preventing corrosion and the social costs of corrosion, or about the fact that most compounds of transition metals are coloured. This, however, makes it difficult to classify environmental phenomena in a way that will give easy retrieval.

Put down your test tube and take to the streets!

Postscript

For more than 10 years the journal, *Education in Chemistry*, has had a regular column pointing out the rich chemical potential of most local environments. This journal is sent free-of-charge to all UK secondary schools by the Royal Society of Chemistry. Past copies should be available in libraries in universities or the new network of Science Learning Centres. Table 1 (overleaf) summarises the content of those columns. Some of the ideas, together with specific examples, are explored in a recent book (Borrows, 2004) and also on a website developed by the Science Learning Centre London (SLC London, 2005).

References

- Borrows, P. (1984) The Pimlico chemistry trail. *School Science Review*, **66**(235), 221–233.
Borrows, P. (2004) Chemistry trails. In *Learning science outside the classroom*, ed. Braund, M. and Reiss, M. pp. 151–168. London: RoutledgeFalmer.

Websites

- For air pollution data:
<http://www.stanger.co.uk/siteinfo/>
For river pollution data:
www.environment-agency.gov.uk/maps
Science Learning Centre, London (2005):
www.sciencelearningcentres.org.uk/london

Peter Borrows is Director of the CLEAPSS School Science Service.

Table 1 The *Education in Chemistry* series on chemistry in the local environment

Year	Vol.	Month	Issue	Page	Main chemical ideas covered
1994	31	January	1	7	Concrete, thermal decomposition, neutralisation, effect of heat on carbonates, stalactites.
		May	3	63	Granite and sandstone, effect of acid on carbonates, expansion of water on freezing, crystals.
		September	5	118	Bricks and mortar, colours of transition metal compounds, acid/base reactions, crystals on bricks.
1995	32	January	1	6	Limestone. Portland Stone, marble, fossils, acid rain, effect of acid on carbonates, lichens.
		May	3	62	Rusting of iron, expansion of iron when rusted, electrolytic nature of corrosion, protection from corrosion, differential aeration, galvanising, sacrificial anode.
		September	5	120	Zinc, lead, copper, brass, bronze, aluminium, chromium, sodium, mercury, tungsten, corrosion of metals, crystals of metals, colours of transition metal compounds.
1996	33	January	1	7	Glass, colours of glass, colours of transition metal compounds, recycling, pebbles on the beach.
		May	3	63	Salt, the salt industry now and in history, chemical place names, lowered freezing point of salt solutions.
		September	5	119	Plant pigments, colours of autumn leaves, chlorophyll and decomposition products, chromatography, indicators from plant materials.
1997	34	January	1	8	Statues, plaques and graves of famous chemists, chemical events, generating hydrogen 200 years ago to fill balloons, action of acid on metals.
		May	3	63	Chemical place names, town and streets, sites of chemical interest.
		September	5	120	Swimming pools, chlorine and substitutes, pH indicators, neutralisation.
1998	35	January	1	8	Hydrocarbons, alkanes, greenhouse gases, methane, gas cylinders, petrol stations, tars, candles, polythene.
		May	3	63	On the beach, crystals of salt, hydrogen bonding, sulfides, silicates, pollution.
		September	5	118	Plastics, polymers, recycling symbols.
1999	36	January	1	6	Air pollution, sulfur dioxide, nitrogen oxides, carbon monoxide, ozone, particulates, air quality monitoring.
		May	3	63	Water, rivers, streams, reservoirs, hard water, springs, bottled water, Epsom salts, chlorine, ozone, copper sulfate, alum, pH, action of acid on metals.
		September	5	118	At the supermarket: acids, alkalis, salts, elements and compounds.

Table 1 The *Education in Chemistry* series (continued)

Year	Vol.	Month	Issue	Page	Main chemical ideas covered
2000	37	January	1	10	Fireworks, oxidation and reduction, flame colours, rates of reaction.
		May	3	66	Transporting chemicals, Hazchem warning panels, emergency procedures.
		September	5	118	Light sources: incandescent and fluorescent lamps, tungsten halogens, mercury, sodium, argon, neon, candles.
2001	38	January	1	23	Tea, coffee and fizzy drinks, caffeine, non-sugar sweeteners, filtration, rates of reaction, solubility of gases in water
		May	3	63	Food: E-numbers, acids, bases, salts, elements.
		September	5	121	Bread, butter, jam; ethanol, colloids, atoms and molecules; catalysis.
2002	39	January	1	14	Air, argon, noble gases, carbon dioxide, greenhouse effect, atoms and molecules, radioactivity, radon.
		May	3	64	Lead metal, pewter, solder. Uses related to properties – easy to extract from ores, chemically unreactive, malleable, low melting point.
		September	5	119	Lead compounds, car batteries, yellow lines on the road, colours of transition metal compounds, lead-based pigments (oxides, carbonate, sulfide), lead compounds in glass, toxicity, action of acids on lead.
2003	40	January	1	9	Gases and their uses. Air, helium, carbon dioxide, ethene, ethyne, coal gas, natural gas, mole calculations, radioactivity
		May	3	62	Smelly gases: chlorine, ozone, ammonia, hydrogen sulfide, nitrogen oxides, acid rain, diffusion.
		September	5	119	Jewellery metals. Gold, silver, platinum, titanium. Brass, nickel, corrosion and tarnishing of these metals.
2004	41	March	2	33	Jewels: rubies, sapphire, amethyst, quartz, pearls, action of acid on carbonates, crystal lattices, colloids.
		September	5	119	Rare earths (lanthanides): self-cleaning ovens, catalysis, cigarette lighter flints, reactions of metals, neodymium magnets, fluorescers.
2005	42	January	1	8	Radioactivity: potassium, 'lo-salt', fly ash, garden soil, half-life.
		May	3	63	Radioactivity: uranium in the garden, radon, thorium, gas mantles, smoke alarms
		September	5	119	(Coloured) compounds of iron, pigments, glass, oxidation, blood, colour of clay and of bricks, salt crystals, E-numbers.
2006	43	In press			Other compounds of iron. Iron(II) sulfate – historical production and uses.