

The chemistry of limestone: teacher's notes

Level

This activity is most appropriate for students aged 14-16 to illustrate chemical reactions and useful materials made from rocks.

Topic

This activity illustrates some of the simple chemical reactions of **limestone** (calcium carbonate, CaCO_3) and **lime** (calcium oxide, CaO).

Description

The activity is suitable as a class practical or as a demonstration. The students (or, less preferably, the teacher) heat **limestone** (mainly calcium carbonate) to form **lime** (calcium oxide) and note the differences between the reactions of limestone and of lime with water, acids and carbon dioxide.

Context

Students should know about differentiation of materials, eg rocks, on the basis of physical properties, and the activity assumes that some work has already been done on the physical examination of rocks. Students should know that carbon dioxide is a gas and have simple ideas about reversible and irreversible changes. They should be aware of simple properties of acids, alkalis and indicators.

The activity concentrates almost exclusively on chemistry, but there are also important potential links with fossils and evolution.

Teaching points

The chemical and physical properties of limestone, especially when reasonably pure, make it highly sought after for hundreds of everyday uses. With salt and coal, it formed the main feedstock for the chemical industry until about 1914. It is still important today as shown by the wide range of uses in Tables 1 - 4 (Appendix). The chemistry is relatively straightforward and can be used to illustrate many types of simple reactions and properties. It is then possible to relate these to industrial and domestic applications (see *Limestone in everyday life*).

Timing

It should be possible to carry out the activity either as a class practical or as a demonstration within a teaching period of about one hour.

Apparatus

Each student (or group) will need:

- eye protection
- Bunsen burner, tripod and gauze
- heatproof mat
- tongs
- 3 test-tubes
- test-tube rack

- dropping pipette
- drinking straw

Chemicals

Each student (or group) will need:

- a few small lumps of limestone (each about 1cm^3) (limestone includes chalk). Marble chips will do if no local source of limestone is available
- deionised / distilled water
- Universal Indicator solution and colour chart

Safety notes

- Wear eye protection.
- Take care when heating as the lumps will become very hot.
- Calcium oxide (lime), the material formed when the lumps are heated, is corrosive. It causes burns and is irritating to eyes, skin and the respiratory system. The reaction of calcium oxide with water is vigorous and exothermic.
- It is the responsibility of the teacher to carry out an appropriate risk assessment.

The activity

For fuller details of the experiment, see the student's material.

Students take about half a dozen small (about 1cm^3) lumps of limestone. They examine the stone and describe briefly its colour, texture and any other notable features such as fossils. The colour of a piece of limestone may be misleading. For example coarse brown **limestones** may be wrongly described as **sandstones**. (Limestone comes in almost every imaginable hue – from white, through yellows, reds, oranges, blues, purples, olives to browns and black.) These colour variations are almost all due to iron content. Some of the darker colour may be due to carbon or possibly manganese. (If no local source of limestone is available, marble chips, available from the prep room, will do.)

Students heat a couple of lumps on a tripod and gauze with a roaring Bunsen flame for 15 minutes. If possible darken the room briefly to allow students to note what happens when the flame is trained directly on the lumps. It may be possible to see the lumps glowing – this is the origin of the term 'limelight'.

After allowing the lumps to cool, students compare the heated lumps with unheated ones.

Lumps that have been heated:

- may appear whiter than the unheated ones
- should crumble more easily than the unheated ones
- will react exothermically when a few drops of water are added
- will show an alkaline pH

Blowing through a straw into the clear solution formed by reacting the heated lumps with water will turn the solution cloudy.

Note. If it is necessary to spread the practical work over two teaching periods, teachers should be aware that, in the intervening period, the lime (calcium oxide) produced by heating the limestone may combine with carbon dioxide from the air to

re-form calcium carbonate, thus reducing its reactivity very significantly. It would be worth making some fresh lime just before the second lesson.

The chemistry of the reactions is as follows:

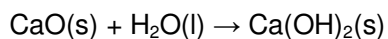
Heating the limestone (calcium carbonate) drives off carbon dioxide gas leaving behind lime, the base calcium oxide.



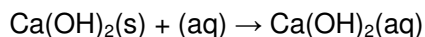
The lime is white and will have a more crumbly texture than the original limestone.

Calcium carbonate does not react with water.

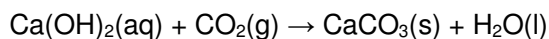
Adding water to the lime produces slaked lime (calcium hydroxide) in an exothermic reaction.



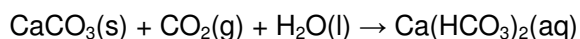
Some of the calcium hydroxide dissolves in the water producing an alkaline solution called limewater.



On blowing into this solution through a straw, the calcium hydroxide solution reacts with the carbon dioxide in exhaled breath to form a cloudy precipitate of calcium carbonate (this is the basis of the limewater test for carbon dioxide). In effect, we have regenerated the original limestone.



Continuing to blow through the straw for some time will result in the calcium carbonate precipitate re-dissolving as soluble calcium hydrogencarbonate.



Appendix: limestone data for Great Britain and Northern Ireland

Use	Quantity / kt
Construction	
Roadstone coated	9175*
Roadstone uncoated	22 481*
Railway ballast	99*
Concrete	15 309*
Other	29 262* (a)
Cement	9831
Building stone	301 (b)
Asphalt filler / mine dust	216
Building lime	460
Industrial	
Agricultural / horticultural	795
Iron & steel	3239 (c)
Specialist fillers	875 (d)
Soda ash	1000 (e)
Sugar refining	250 (e)
Glass	203
Other lime n.e.s.	139
Other uses n.e.s.	666
Total	94 547 (f)

Table 1 The uses of limestone in Great Britain and Northern Ireland(1999)

Notes on Table 1

(a) mainly foundation and fill

(b) *ie* architectural, walling, dimension stone

(c) mainly iron- and steel-making flux

(d) powders + 'whittings' used in animal feeds, polymers (plastics, rubber) paint, paper, pharmaceuticals

(e) estimated

(f) in addition about 1.8 Mt of dolomite were used for industrial purposes (especially furnace linings and production of magnesium compounds, notably magnesia)

* construction aggregates total = 76 326

n.e.s. not elsewhere specified

Sources: British Geological Survey, Minerals Year Book; Office for National Statistics; National Stone Centre

Country	Quantity / kt
England	72 820
Wales	17 220
Scotland	1507
N.Ireland	4219 (a)
Total	98 766

Table 2 Production of limestone in Great Britain and Northern Ireland (1999) by country

Notes on Table 2

(a) almost all for aggregates – figure includes hard chalk

Producing area (a)	Quantity / kt
Derbyshire (inc. Peak National Park)	19 240
Somerset	11 550
N.Yorkshire	7528
Clwyd	7269
Mid Glamorgan	5076
Lancashire	5072
Avon	4948
Durham	4401
Cumbria	4389
Leicestershire	3419

Table 3 Production of limestone in Great Britain and Northern Ireland (1999) by producing area

Notes on Table 3

(a) N.B. recent data in some cases published for counties which were reorganised in 1990s

N.B. all figures in Tables 1 – 3 (except N.Ireland) are for limestone excluding chalk

Region	Quantity / kt
England	9667
Of which: South East Region	4144
Yorkshire/Humber Region	3268
Of which: cement	6345
construction	1021
Misc. uses (inc. fillers)	1701

Table 4 Production and uses of chalk in England (1999)