Limestone to Quicklime

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Limestone is made of fossil shells that were almost pure calcium carbonate. Modern shells are easy to find. We will use them as our source of *calcium carbonate* ($CaCO_3$).

To convert calcium carbonate (limestone) to *calcium oxide* (*quicklime*) the carbonate must be heated to close to 840 °C.
To do this at home we will use a small charcoal fire: hoping that's hot enough. The shells will be put in a small steel can. *This demonstration is not usually done in schools because Bunsen Burners will not heat calcium carbonate to a high enough temperature.*

The chemical reactions are ….

$$CaCO_3 \rightarrow CaO + CO_2$$

$$3Fe + 2O_2 \rightarrow Fe_3O_4$$

The can has a mass of 28 grams and the mass of the shells is 100 grams.

A small outdoor stove has been loaded with charcoal and a can of shells is buried in the pile.

More charcoal is added before the fire becomes too hot.

The charcoal is burning nicely after 20 minutes.

After another 30 minutes the fire is a bright orange inside.

The fire is left to burn out and the can is removed while still warm.

The can is now covered inside and out with black iron oxide and has a total mass of 33 grams. The shells now weigh just 73 grams.

Theory

1 The atomic weights of Ca, C and O are 40, 12 and 16 respectively. The formula weight of $CaCO_3$ is 100. For complete conversion …

$$CaCO_3 \rightarrow CaO + CO_2$$

… so 100 grams of shells would become 54 grams of *quicklime* (CaO).

Because the mass of the shells has been reduced to only 73 grams, about half of the $CaCO_3$ has been converted to CaO.

2 The atomic weights of iron and oxygen are 56 and 16. The formula weight of Fe_3O_4 is (3x56+4x16) = 232. For complete conversion …

$$3Fe + 2O_2 \rightarrow Fe_3O_4$$

168 g of Fe → 232 g of Fe_3O_4 … so … 28 g of Fe → '*x*' g of Fe_3O_4

$$x = 232x28/168$$

$$= 39 \text{ g}$$

The mass of the can has increased to 33 grams so the conversion to iron oxide is incomplete. The can has not completely disintegrated into black flakes, which in this case is a convenient outcome.

The 25 or so grams of CaO that has been made is sufficient for the next part of the process: converting CaO to $Ca(OH)_2$ by adding water. $Ca(OH)_2$ was once called *slaked lime* and now has the modern chemical name of *calcium hydroxide*.

A Planned second attempt

A higher temperature (and/or perhaps more time) is needed to drive both reactions to completion and verify that complete conversions do give the expected weights of the products. A larger fire that burns hotter for longer is needed. It may also be necessary to force air through this larger fire to increase the rate of burning.