The Persistence of the Candle-and-Cylinder Misconception W

James P. Birk[1]
Department of Chemistry and Biochemistry, Arizona State University, Tempe, AZ 85287; *jbirk@asu.edu*

Anton E. Lawson[1]
Department of Biology, Arizona State University, Tempe, AZ 85287; *Lawson@asu.edu*

Place a candle in a container of water, held upright by a small piece of clay. Ignite the candle and lower a beaker over the candle to trap a sample of air. After a short time, the candle flame is extinguished and the water rises into the beaker. The change in water level corresponds (*sometimes*) to a 21% change in the enclosed gas volume, presumably showing that the consumed oxygen occupied 21% of the air in the beaker. This experiment has been around for some time and has often mistakenly been used as a method for determining the percentage of oxygen in air. Variations of the experiment involve the use of burning matches, paper, "inextinguishable" birthday candles, or alcohol in place of the candle, and a graduated cylinder, test tube, glass jar (or tumbler), or round-bottom flask in place of the beaker.

But, even when these experiments result in a 21% change in gas volume, do they measure the percentage of oxygen in air? Clearly they do not, but the misconception that this is a valid method for measuring the oxygen content of air seems to persist nevertheless. A number of out-of-date chemistry texts describe the candle-and-cylinder demonstration, which also purports to show that oxygen comprises 21% of normal air (e.g., *1–4*). Two recent publications have resuscitated this misconception (*5*, *6*). The invalidity of these experiments has been demonstrated previously (*7*, *8*), and we present more evidence here to show that the oxygen is not completely consumed in these experiments. Elsewhere, we have discussed the instructional and scientific consequences of designing experiments and modifying experimental design in order to obtain a predetermined result that matches an existing conception, which may actually be a misconception (*9*).

The candle-and-cylinder demonstration is invalid because the rapid water rise can be almost completely attributed to the expansion and escape of heated air, rather than to the "consumption" of oxygen. That is, the burning candle causes the surrounding air to expand and escape. Thus, when the candle goes out, the expanded air that remains trapped in the cylinder cools and contracts. This lowers the internal air pressure so that the relatively greater external air pressure pushes down on the external water's surface, pushing water up into the cylinder until the internal and external pressures equalize—to a level that occasionally and coincidentally turns out to be roughly 21% of the initial cylinder volume. Close examination of this experiment in progress shows that there is often an initial expansion of the gas volume, due to heating, and that the contraction of gas volume does not occur until the flame begins to expire. If an appreciable fraction of the oxygen were consumed during the experiment, we would expect to see a gradual contraction of the gas volume as the oxygen is consumed.

Evidence that supports the above explanation and allows us to reject the oxygen-consumption hypothesis also comes from a number of experiments described by Peckham (*7*). We add to these a test of the experiments described by Caplan et al. (*5*). The usual apparatus is preassembled with an unlit candle having a match attached to the wick. The water level in the cylinder is raised part way into the beaker by removing air with a syringe. The raised water level insures that no water will escape from the bottom of the beaker when the match and candle are lit. The candle (and match) are ignited by focusing sunlight with a magnifying glass. After combustion starts, the heated and expanded air pushes the water level down inside the beaker, but not far enough to allow any gas to escape out the bottom. Then when the match and candle go out and the expanded air cools to the ambient temperature, the final water level inside the beaker provides information about what happened during and after combustion.

Suppose we assume that the candle wax consists of the hydrocarbon pentacosane ($C_{25}H_{52}$). Then the balanced equation for combustion is

$$C_{25}H_{52} + 38O_2 \rightarrow 25CO_2 + 26H_2O$$

Based on this equation and possible fates of the reaction products, we can make some predictions about the expected volume change:

1. If all the O_2 is converted to CO_2 and H_2O and if both products remain in the gaseous state, the final water level should be *lower* than the initial level by about 7.2% (38 gas molecules [21.0%] $\rightarrow$ 51 gas molecules [28.2%]).
2. If all the O_2 is converted to CO_2 and H_2O and if the CO_2 quickly dissolves into the water but the H_2O remains in the gaseous state, the water should rise about 6.6% above the initial level (38 gas molecules [21.0%] $\rightarrow$ 26 gas molecules [14.4%]).
3. If all the O_2 is converted to CO_2 and H_2O and if the CO_2 quickly dissolves and the H_2O quickly condenses, the water level should rise 21% above the initial level (38 gas molecules [21.0%] $\rightarrow$ 0 gas molecules [0%]).
4. If all the O_2 is converted to CO_2 and H_2O and if the CO_2 remains in the gaseous state but the H_2O quickly condenses, the water level should rise about 7.2% (38 gas molecules [21.0%] $\rightarrow$ 25 gas molecules [13.8%]).
5. If combustion is incomplete in the sense that something less than the presumed 21% O_2 is converted to CO_2 and H_2O, then each of these predicted water level changes (alternatives 1–4) should be less by some undetermined amount.

The experiment was replicated several times, giving reasonably consistent results. After the match and candle go out (about 10 seconds after ignition), droplets quickly appear on the inside of the beaker and the water level rises between 1.7% and 4.1% above the initial level. These results allow us to reject alternatives 1, 2, and 3. Apparently H_2O quickly condenses, but CO_2 does *not* quickly dissolve. Further, it seems clear that combustion of all 21% of the initial O_2 does not occur.

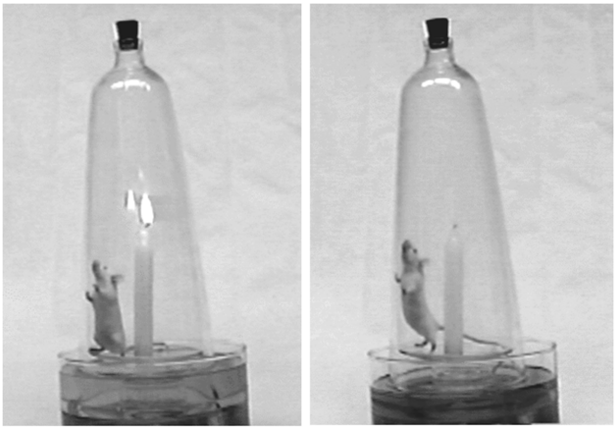

Figure 1. A mouse remains energetic in an atmosphere that no longer supports a candle flame, illustrating that this atmosphere still contains considerable oxygen. Left: Mouse in enclosed container with burning candle. Right: Mouse is still energetic after candle flame expires. Note the change in water levels arising from the cooling of the enclosed air.

This conclusion is further supported by results of an experiment conducted by developers of an instructional unit called *Gases and Airs* (*10*). The experiment, recorded in a film entitled *The Mouse and the Candle*, involves placing a lively mouse under a bell jar with a burning candle. After the candle slowly burns out, the mouse remains as energetic as before. We have also reproduced this experiment (see Fig. 1) and a video is available.[W] Therefore, we can be quite certain that the initial candle-and-cylinder demonstration is not a valid demonstration of the claim that 21% of normal air is oxygen. Caplan et al. (*5*) appear to realize that the initial candle-and-cylinder demonstration is somehow flawed and that several of their subsequent modified experiments are questionable, but they nevertheless fail to reject their initial misconception, continuing to modify their experiment until they claim to have obtained the expected results. Thus, they fail to reject the candle-and-cylinder demonstration as a valid test of the claim that 21% of air is oxygen.

A recent issue of this *Journal* contains another variation on the candle-and-cylinder demonstration, which also claims to show that air contains 21% oxygen (*6*). To set up the experiment, a few drops of ethanol are dropped onto a small wad of cotton that has been placed just below the tip of a hollow injection needle. The needle is inserted through a rubber stopper and connected to a plastic tube, which is closed by a valve. The experiment involves igniting the ethanol and then inserting the burning cotton, the needle, and the rubber stopper into an inverted round-bottom flask. The rubber stopper seals the burning ethanol in the flask. After the ethanol stops burning, the plastic hose is placed in a flask of water. The valve is then opened and water quickly squirts up through the needle into the round-bottom flask until the pressures inside and outside the flask are equalized.

Fang claims that the amount of water that squirts into the flask is a measure of the amount of oxygen in air. On the basis of the results of several replicates of his experiment, he presents a table of highly consistent sample data and calculated results to substantiate the claim that this experiment measures the oxygen content of air to be 21%.

When we tried to replicate Fang's experiment with the same experimental setup, we obtained results that are not consistent with complete consumption of oxygen, but *are* consistent with expulsion of heated air during insertion of the burning ethanol. Specifically, the results depended on the speed with which the stopper was secured in the flask. Volume changes in five experiments were 36.6% (stopper inserted slowly), 42.0% (stopper inserted slowly), 22.3% (stopper inserted quickly), 22.9% (stopper inserted quickly), 17.4% (stopper inserted as quickly as possible). These results are not consistent with the lower internal air pressure in the flask being caused by complete oxygen consumption followed by rapid carbon dioxide dissolution, but are consistent with the hypothesis of escaping hot air. When we pushed the stopper in slowly, more hot air escaped.

We also checked the rate of carbon dioxide dissolution. We used dry ice to fill a 58-mL test tube with carbon dioxide and inverted the test tube into a 250-mL beaker of water. The water in the beaker was replaced by a stream of tap water every 5 minutes. During the course of 1 hour, the water level rose by only 7 mL (12% of the initial volume). This clearly refutes the rapid dissolution of carbon dioxide in water for most of the reported experiments. However, the water spray in Fang's experiments (*6*) would be expected to provide greater surface area and result in faster dissolution.

We tested Fang's experimental setup outside our laboratory building, using a magnifying glass to focus the sun's rays on a match taped to the top of the injection needle. (We chose not to use the potentially explosive ethanol–air mixture for this experiment.) When lit in this way, the match burned for a few seconds. After the air in a 214-mL flask was allowed to cool (by taking the apparatus back inside), water vapor was clearly visible on the inside of the flask. But when the clamp was loosened, water rose only part way up the tube (4 mL or about 2%). Additional experiments in which we did not take the apparatus back into the cooler laboratory but did allow the hot glass to cool to ambient temperature always resulted in expulsion of gas when we loosened the clamp. These results clearly refute the oxygen-consumption hypothesis.

In conclusion, in every experiment that purports to measure 21% oxygen in air by a combustion process, combustion is not consuming all the oxygen. Rather, the heating of air results in partial expulsion of air. The flame is extinguished by local depletion of oxygen and buildup of carbon dioxide, not by complete consumption of oxygen in the remaining air. When the flame expires, the apparatus and enclosed air cool back to room temperature, causing a pressure differential that results in a rise in the water level.

Several valid methods for determining the oxygen content of air have appeared in chemistry textbooks and journals (*11–19*). The method described by Birk et al. (*11*) is particularly unambiguous because it involves rusting steel wool in an enclosed container. As the steel wool rusts, oxygen is incorporated from the gaseous state into the solid state with a corresponding 21% decrease in gas volume. Variations on this experiment have also been published (*12–14*). Other methods include reduction of oxygen by copper in ammoniacal ammonium chloride solution (*15*), by pyrogallol (*16–18*), and by hydrogen gas (*19*).

Acknowledgment

This material is based upon research partially supported by the National Science Foundation under grant DUE 9453610.

Notes

[W] A video of this experiment is available on *JCE Online* at *http://jchemed.chem.wisc.edu/Journal/issues/1999/Jul/abs914.html*.

1. Any opinions, findings, and conclusions or recommendations expressed in this publication are ours and do not necessarily reflect the views of the National Science Foundation.

Literature Cited

1. Cobb, V. *Chemically Active*; Lippincott: New York, 1985; p 45.
2. Mullin, V. L. *Chemistry Experiments for Children*; Dover: New York, 1968; pp 30–31.
3. Farrah, K. H. *Middle School Chemistry*; G. Bell: London, 1965; pp 22–23.
4. Cooper, E. K. *Discovering Chemistry*; Butterworth: London, 1963; pp 118–119.
5. Caplan, J. B.; Gerritsen, H. J.; LeDell, J. S. *Phys.Teach.* 1994, *32*, 310–314.
6. Fang, C. H. *J. Chem. Educ.* 1998, *75*, 58–59.
7. Peckham, G. D. *J. Chem. Educ.* 1993, *70*, 1008–1009.
8. Glanz, J. *J. Chem. Educ.* 1963, *40*, A477.
9. Lawson, A. E.; Lewis, C. M. Jr.; Birk, J. P. *J. Coll. Sci. Teach.*, in press.
10. Elementary Science Study. *A Working Guide to the Elementary Science Study.* Educational Development Center, Inc.: Newton, MA, 1971.
11. Birk, J. P.; McGrath, L; Gunter, S. K. *J. Chem. Educ.* 1981, *58*, 804–805.
12. Birk, J. P.; Kurtz, M. J. *Investigations in Chemistry;* Houghton Mifflin: Boston, 1994; pp 239–240.
13. Meek, E. G. *Sch. Sci. Rev.* 1983, *65*, 90–93.
14. Martins, G. F. *J. Chem. Educ.* 1987, *64*, 809–810.
15. Lincoln, A. T.; Klug, H. P. *J. Chem. Educ.* 1935, *12*, 375–376.
16. Davenport, D. A.; Saba, A. N. *J. Chem. Educ.* 1962, *39*, 617–620.
17. Munro, L. A. *J. Chem. Educ.* 1928, *5*, 741–742.
18. Forster, A. C. *J. Chem. Educ.* 1927, *4*, 638–639.
19. Cooke, G. B. *J. Chem. Educ.* 1929, *6*, 105–106.