

Episode 601: Brownian motion and ideal gases

This episode looks at Brownian motion as evidence for the particulate nature of matter, and the macroscopic gas laws.

Summary

Demonstration and discussion: Brownian motion and what this tells us about air (and other gases). (30 minutes)

Demonstration and student experiments: The gas laws. (60 minutes)

Discussion: Boyle's law – a particle explanation. (20 minutes)

Discussion: Extrapolating to absolute zero. (10 minutes)

Student activity: A computer model of Boyle's law. (20 minutes)

Demonstration + discussion:

Brownian motion and what this tells us about air (and other gases)

A reasonable place to start is by reviewing the evidence that matter is made up of particles (molecules and atoms). (It is probably best to refer to 'particles' in general, and to think of them as spherical; point out that, by this, we mean either atoms or molecules.) Students will have met these ideas at a lower level in chemistry as well as physics. Evidence includes the combination laws of gases, and Brownian motion, which can be demonstrated in the classroom.

According to your students' previous experience, you may wish to demonstrate Brownian motion, the expansion of bromine into a vacuum, and a measurement of the density of air.

TAP 601-1: Brownian motion in a smoke cell

TAP 601-2: Diffusion of bromine

TAP 601-3: The density of air

Brownian motion is evidence not just for the existence of atoms or molecules but also for their movement, which is random. This random motion can be modelled mathematically and leads to a test for the size of the atoms from measured diffusion rates – this was one of Einstein's great papers of 1905.

Question: What is the mass of air in the room? Answering this will require the estimation of the volume of the room, and the use of the density of air. Students are often surprised by the result: perhaps 100 kg, more than the mass of a typical student. If all of the air in the room condensed to form a liquid, it would make a layer perhaps 5 mm deep on the floor.

Now you can introduce a gas as a simple system. In crystalline solids, all the atoms are nicely ordered in an array, making calculations possible. In a gas, the motion is random and again this simplifies calculations as the laws of statistics can be applied. (Disordered solids and liquids are

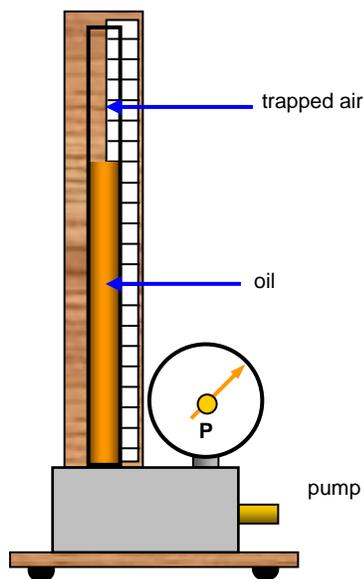
more difficult to treat mathematically, because they are neither well-ordered nor completely disordered.)

Demonstration + student experiments:

The gas laws

Now move on to the gas laws. Strictly speaking it is only necessary to look at Boyle's law ($PV = \text{constant}$ at constant temperature).

You may well have an apparatus specifically designed to show this such as an oil-filled column attached to a pump and pressure gauge. You pump air in to pressurise the oil, which compresses an air space at the top. A series of about 10 P and V readings usually gives a good fit to a straight line when P is plotted against $1/V$.



TAP 601-4: Changes in volume, changes in pressure.

The other laws follow from this and the definition of *thermodynamic* temperature. However, at this stage it is likely that the only idea of temperature students are familiar with is "that which is measured by a thermometer". (This is not a totally silly definition as it relies on the fact that when two bodies of different material, temperature and size are in contact, their temperatures equalise. This is usually referred to as the Zeroth Law of Thermodynamics.)

So it is worth pressing ahead with demonstrations or class experiments of Charles' law (V proportional to T) and/or the pressure law (P proportional to T).

TAP 601-5: Changes in temperature, changes in pressure.

TAP 601-6: Changes in temperature, changes in volume.

Discussion:

Boyle's law – a particle explanation

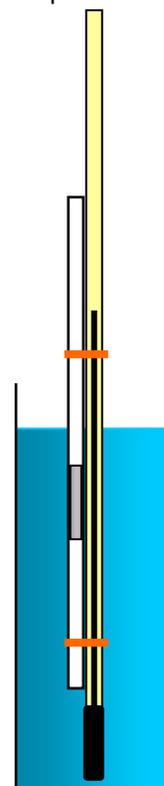
You may need to discuss how Boyle's law can be explained in terms of particles.

TAP 601-7: Boyle's law, density and number of molecules

Discussion:

Extrapolating to absolute zero

Both Charles' law and the pressure law lead to the concept of an absolute zero of temperature. Note that Absolute Zero (0 K) is the temperature at which the energy of the particles of a material has its



(resourcefulphysics.org)

Charles' law apparatus

minimum value; this is not zero, as the particles have so-called zero point energy due to quantum effects, which cannot be removed. They still vibrate.

Remember real gases turn into liquids and solids before absolute zero is reached. However extrapolating back often gives reasonable values.

TAP 601-8: Changing pressure and volume by changing temperature.

TAP 602: Ideal gases and absolute zero

Student activity:

A computer model of Boyle's law

If suitable apparatus for demonstrating Boyle's law is not available it can be simulated using computer packages or applets but this is a poor relation to the real experiment and should only be used to illustrate the effect, not to demonstrate it.

<http://chem.salve.edu/chemistry/boyle.asp>

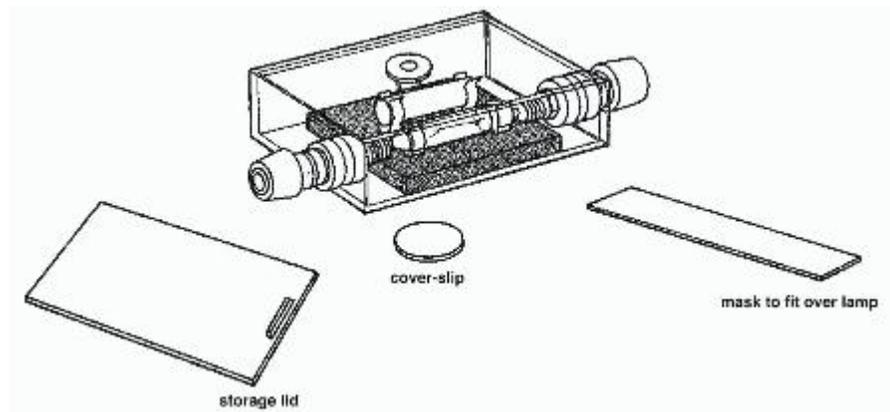
TAP 601-1: Brownian motion in a smoke cell

This is a 'classic' experiment that gives strong circumstantial evidence for the particulate nature of air.

You will need:

- ✓ Smoke cell, incorporating a light source and lens (Whitley Bay pattern)
- ✓ Microscope, low power (e.g. x10 objective, x 10 eyepieces) and large aperture
- ✓ Power supply, 0 to 12 V dc
- ✓ Microscope cover-slip
- ✓ Smoke source (e.g. paper drinking straw)

Set up



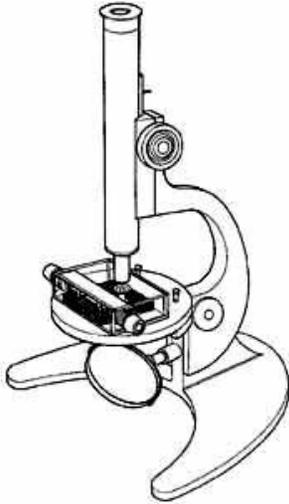
The smoke can come from a piece of burning cord using a dropping pipette or a burning straw (preferably paper). The straw should burn at the top and then be extinguished. The bottom end of the straw should poke into the plastic smoke container.

The cell may need to be cleaned if a waxy or plastic straw is used.

Remove the glass cell from the assembly to clean it. Afterward, push it fully back into the assembly. It may help to wet the outside of the glass tube. You will find it helpful to clean the glass cell after every five to ten fillings to obtain the best results; otherwise the light intensity is reduced.

The cell is illuminated from one side to make the smoke particles visible under the microscope. This is called dark ground illumination. A small piece of black card prevents stray light from the lamp reaching the eye. The lamp is placed below the level of the glass rod in order to minimise convection.

What to do



1. Fill the cell with smoke using a dropping pipette and cover it with a glass cover-slip. This will reduce the rate of loss of smoke from the cell
2. Place the cell on the microscope stage, fit the mask and connect to a 12 V power supply.
3. Start with the objective lens of the microscope near the cover-slip. While looking through the microscope, slowly adjust the focus, moving the objective lens away from the cover-slip, until you see tiny dots of light.
4. Watch the particles carefully. Note what you see.

Practical advice

As this is such an important experiment - one of the few to show the 'graininess' of nature and to give strong support to the idea that gas molecules are in constant motion - students should be given plenty of time to set it up and see it clearly.

We know what the students are supposed to see. They may not. Consequently, the students may not 'see' what we expect. We expect them to observe jiggling points of light. The vertical component of the motion causes the bright points to go out of focus and to disappear. This will not be obvious to every student. The points of light may also have a drift velocity but we know that this observation is unimportant. The students don't know that the drift (due to large scale convection effects amongst other causes) is unimportant, and so this may become their major observation. A 'prepared' mind helps the scientist to see.

Once the students 'know' what to look for, it is useful to repeat the experiment – they'll see the expected effect second time around.

The bright specks of light do not bounce into each other before changing direction. Why?

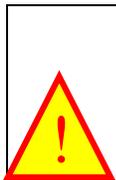
Ask students to note down an explanation of the observations. Discuss what everyone thinks is going on and then describe (or elicit through questioning) the kinetic explanation – that smoke particles are observed as points of light, and their jiggling is due to collisions with much smaller and invisible air particles. Students should realise that an invisible movement explains an observed movement. It becomes somewhat circular when Brownian motion is (incorrectly) given as evidence of the particulate nature of air – it is circumstantial at best.

"How big are air molecules?" "Well, smaller than the smallest specks of ash that makes up smoke!" This may lead to guesstimating.

1 mm across?, 0.1 mm?, 0.01 mm? Or even smaller? How many smoke particles could you park side by side along the edge of a postcard (15 cm long)? This is a chance to make an order of magnitude guess on whether they think that there would be 100, 1000, 10000, 100000, working in powers of 10. And now guess how many molecules you could park side by side on each smoke speck. How many is that along the edge of the postcard? Molecules must be very small indeed and atoms even smaller. Common sense tells us that.

Alternative approaches

If a camera is available for the microscope, you can demonstrate this experiment quickly to the whole class by following the camera instructions. However, 'seeing for yourself' has much to commend it in this case.



Safety

Teachers must ensure that light from the Sun is not reflected up through the microscope. (Once the cell is in position, the stage aperture is covered, removing the hazard.)

Brownian motion: facts and myths

Robert Brown is correctly referred to as having observed the jittering motion of small particles, but

1. he wasn't the first to record the observation,
and
2. he did NOT observe the motion of actual pollen grains.

How many text books etc continue to hand on these mistakes?

The title of Browns paper was "A Brief Account of Microscopical Observations ... on the Particles CONTAINED in the Pollen of Plants". Pollen itself is too large (and hence has too much mass) to be small enough to be buffeted significantly by water molecules etc. [The most recent reference pointing this out: *Nature*, 10 March 2005 p 137]

The first recorded observation of what we now call Brownian motion was made in 1785 by Jan Ingenhauz using charcoal dust. [Ref: *Nature*, 7 June 2001 p 641]

Having used particles derived from living matter, Brown had to try several other inanimate substances to convince himself that the motion he observed was not something to do with a 'life force', but a property of all microscopic matter. This 'systematic investigation' is what won for Brown the accolade of having the jittering motion named after him, work that Ingenhauz didn't need to do.

Today's research into nano technology now routinely fabricates nanoparticles. Controlling them suspended in liquids is quite a task. One method is to use a direct current controlled by a feedback system to cancel out the Brownian Motion. The position of the 20 nm polystyrene spheres is monitored by a fluorescence microscope and the voltage across the solution altered accordingly. So far nano-particles have been confined to within 1 micron. Alternatively, the path of the particle can be manipulated by suitable changes of the applied voltage. [Ref: *Nature*, 10 March 2005 p 156.]

Even before the recent advent of nano-technology, Einstein's 1905 paper on Brownian Motion is his most cited paper (i.e. more than for Special Relativity or his work on photons). It is used by scientists working on such varied topics as aerosol particles ("pollution"), the properties of milk, paints, granular media (powders) and semiconductors. [Ref *Nature*, 20 January 2005 p 216]

External references

This activity is taken from <http://www.practicalphysics.org>

'Brownian motion: facts and myths' is taken from CAPT (Connecting Advancing Physics Teachers) e-mail support by Rick Marshall.

TAP 601-2: Diffusion of bromine

A classic demonstration that is explained satisfactorily by a particulate model of gases. The students first observe relatively sluggish diffusion of bromine into air. This is followed by the much more rapid diffusion (i.e. expansion) into a vacuum

Almost all gases are transparent and so invisible. The following experiment uses bromine, which is brown; however, this is a dangerous, corrosive gas and great care must be taken.

You will need

- ✓ Vaseline
- ✓ one small brush for cleaning stopcocks
- ✓ vacuum pump
- ✓ pliers
- ✓ 500 ml 1 M (25%) sodium thiosulphate solution
- ✓ two bromine diffusion tubes with matched accessories, including the following:
 - ✓ two large bore 8 mm stopcocks
 - ✓ two lengths of rubber tubing (length 125 mm, internal diameter 12 mm)
 - ✓ two borosilicate glass cap tubes, to hold ampoules
 - ✓ two rubber bung no 25 with 11 / 12 mm hole
 - ✓ two bromine ampoules

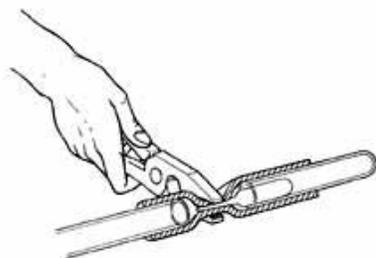
	<p style="text-align: center;">Safety</p> <p style="text-align: center;">Wear Safety spectacles</p> <p>Bromine gas is very toxic and must not be inhaled. The liquid is also corrosive. The teacher must have 500 ml of 1 M sodium thiosulfate solution in a wide beaker so that a hand with a splash of bromine liquid on it can be plunged in immediately. The rubber tubing must fit tightly otherwise the demonstration should be done in a fume cupboard.</p> <p>Use safety screens for the vacuum demonstration. Ensure the technician has been instructed in safe disposal and cleaning.</p>
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What to do:

Diffusion into air

Initially the bromine capsule is placed in the glass tube and the tap is closed. The diffusion tube is full of air.

1. Slide the capsule down from the glass tube into the rubber tubing.
2. Using square pliers, not side-cutters, break the capsule by squashing the rubber tubing. This will release liquid bromine and it will run towards the tap.



3. Watch carefully, as the tap is opened so that bromine moves into the diffusion tube.
4. Measure how far the average bromine molecule (average density of "brownness") moves in about 20 minutes. Calculate the average speed of a bromine molecule.

Diffusion into a vacuum

1. Initially only the bung and tap are attached to the diffusion tube.
2. With the tap open, use a vacuum pump to extract air from the diffusion tube.
3. Once the air has been extracted from the tube, close the tap then remove the vacuum pump.
4. Using rubber tubing, attach to the tap the glass tube with a bromine capsule inside.
5. Slide the capsule down from the glass tube into the rubber tubing.
6. Using pliers, break the capsule by squashing the rubber tubing. This will release liquid bromine and it will run towards the tap.
7. Watch carefully as the tap is opened so that bromine moves into the diffusion tube. It happens almost instantaneously!

Practical advice

The crucial point to these demonstrations is that the 'actual' speed of bromine molecules does not change from one demonstration to the next. Careful questioning of the class will highlight many common misunderstandings. These include the idea that the bromine is being 'sucked' into the tube in the vacuum demonstration.

Alternative approaches

It may be best to discuss precisely what has been observed without any theoretical inferences. This can be followed up by developing a theoretical explanation of the observations based on the particulate picture of a gas.

Teacher and technician note

Bromine gas is very toxic and must not be inhaled. The liquid is also corrosive. The teacher must have 500 ml of 1 M sodium thiosulfate solution in a wide beaker so that a hand with a splash of bromine liquid on it can be plunged in immediately.

The main diffusion tube is a closed glass tube (45 cm long, 5 cm in diameter) with only one opening to a side tube. There is, therefore, no danger of an accident releasing bromine to the pump when diffusion into a vacuum is done. A rubber bung fits into the side tube and carries the glass tube of the stopcock. The glass tube from the stopcock extends through the rubber bung, thus ensuring that only bromine vapour and not liquid comes into contact with the bung. In any case, the bung can and should be replaced, after a few days' use.

The tube in the rubber bung leads to a stopcock with large bore. This should be of good quality, such as Interkey, with bore at least 8 mm. The tap of the stopcock must be spring-held for safety (but the stopcock may be of ordinary quality, not the special high-vacuum quality). The glass tube that leads out on the other side of the stopcock is joined to a closed glass cap by a short section of rubber tube, in which the bromine capsule is to be broken. That rubber tubing must have a fairly thin wall so that it can be squeezed with pliers to crush the capsule.

With this arrangement, the breaking of the capsule to release bromine is done separately, before the stopcock is opened to admit bromine to the main tube. This enables the experimenter to concentrate on the crushing of the capsule first and then pay full attention to the actual demonstration.

Note

The rubber tubing should not be very short; otherwise there is a danger of pulling it off the glass tube when squeezing it with pliers. Rubber tubing must, of course, have a bore large enough to let the capsule slide into it. The tube belonging to the stopcock and the cap tube must be still larger, so that the rubber tube fits tightly on them.)

If the apparatus meets the above criteria, the experiment may be done in the open laboratory using safety screens. If there is any doubt, the demonstration should be done in a fume cupboard.

External references

This activity is taken from <http://www.practicalphysics.org> and is based on Advancing Physics chapter 13 activity 160D

TAP 601-3: The density of air

Background

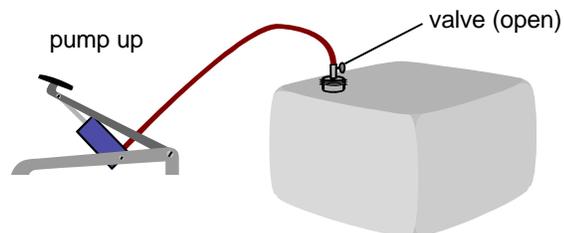
The density of air is so small compared with solids or liquids that we may sometimes consider it insignificant. Yet when a stiff wind blows, we know that air has mass. And there is a lot of it, for example in a room or in a hot-air balloon. This simple experiment allows you to find its density.

You will need

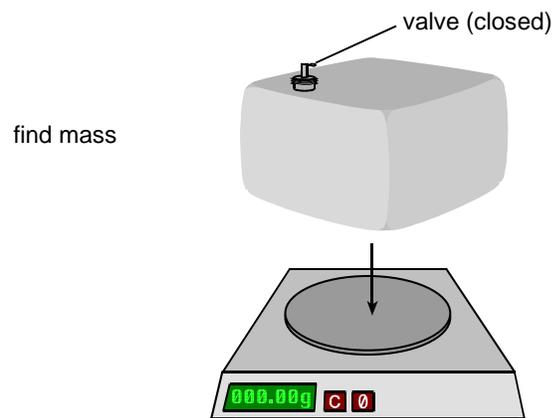
- ✓ large polythene container with tap (e.g. collapsible water container for camping)
- ✓ foot-pump
- ✓ bucket or trough
- ✓ water
- ✓ rubber connecting tube
- ✓ measuring cylinder 100 cm³
- ✓ mass balance, electronic 0 – 1 kg \pm 1 g

What to do

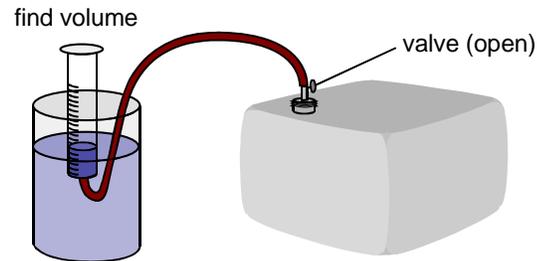
1. Using the foot-pump, fill the polythene container with air until it is hard.



2. Using a balance, find the mass of the container plus air.



3. Fill the measuring cylinder with water and invert it over a bucket of water, so that the cylinder remains full of water.
4. Release excess air from the container until it is again at atmospheric pressure, finding the volume released. This can be done by bubbling the air through the connecting tube and into the measuring cylinder. If the cylinder becomes completely filled, use repeat fills to find the total volume of air released.



5. Find the mass of the container plus remaining air. The difference between this mass and the mass found earlier is of course the mass of the released air.
6. Use the mass and volume of air released to calculate its density.
7. Explain why it was necessary to release the air into the measuring cylinder, rather than simply finding the volume of the plastic container.

You have shown

1. That the density of air is about 1.2 kg m^{-3} or 0.012 g cm^{-3} .

Practical advice

This is a quick and easy experiment, which gives good results. If students have not previously seen an experiment to determine the density of air, then it is valuable for them to see one.

Alternative approaches

You might use a vacuum pump to evacuate a 500 ml round-bottomed flask. The flask should have a one-hole bung and pressure tubing capable of being sealed with a Hoffman clip. Find the mass of the flask plus air before and after evacuation. By opening the Hoffman clip while the tubing is submerged in water, the volume of evacuated air can be found. It is the same as the volume of water entering the flask.

You will need a safety screen for this experiment. The experiment using the plastic container described above is preferred because it leads to a result more simply, and can be done by a student.

Social and human context

Through mining experience, it had been known since ancient times that no pump could lift water higher than 32 feet. Only in the seventeenth century, starting with Galileo's thinking, was this linked to the fact that air has weight and exerts pressure. Thus the mental picture was transformed from sucking water up to using atmospheric pressure to support a column of water.

External reference

This activity is taken from Advancing Physics chapter 13, 10E

TAP 601- 4: Changes in volume, changes in pressure

Common behaviour

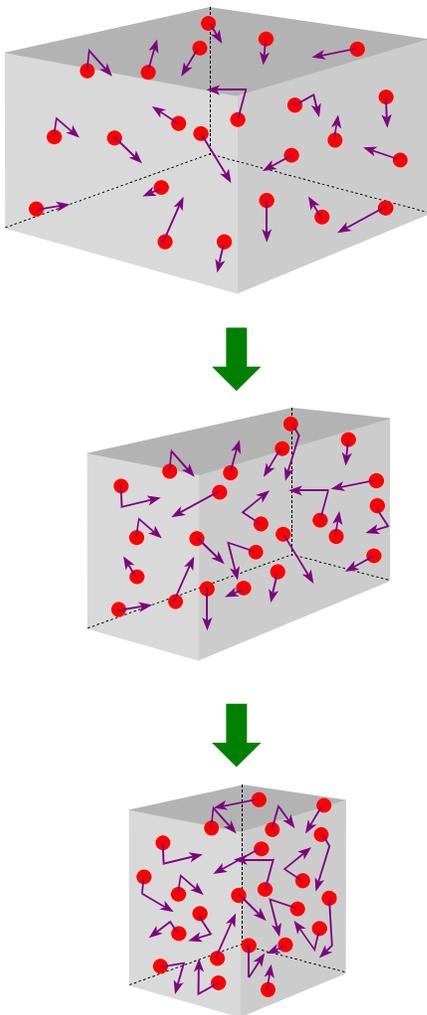
Gases are remarkable because they are all so similar. Solids vary considerably because their particles are tightly bound together, and the details of the bonding affect the properties of the material. Gas particles are not tightly bound and spend most of their time not interacting with other particles, so are much simpler. Once you know the mass and speed of the particles that make up the gas, you know enough to be able to predict some macroscopic properties of the gas.

Here you look at how packing more and more gas into a given volume affects the pressure. Precisely because all gases behave in very similar ways, you do not need to worry about which gas you use.

You will need

- ✓ Boyle's law apparatus

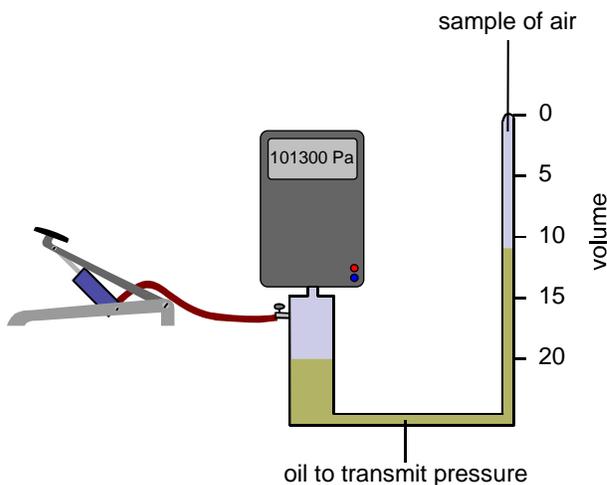
Thinking about the measurements



You need to measure how the volume of a fixed number of particles affects the pressure. Squeezing these particles into a smaller and smaller volume results in more and more collisions with the walls, giving a higher pressure. However, you will only get a true relationship between pressure and volume if the number of particles stays the same: you need to make sure that no molecules escape. Rough and ready results can be obtained by using syringes, but these leak, so more precise ways have evolved – sealing off a volume of gas behind a liquid makes for a good seal. It is the measurement of volume that turns out to be the difficult one to get right. You may be able to set up a slicker arrangement using automated data capture, but you will need to take great care to measure the volume, ensuring the equivalent of a leak-proof syringe, where you can measure the position of the plunger or piston.

Compressing a gas will warm it up and vice versa, so after changing the volume leave sufficient time for the temperature to return to its original value.

A traditional solution



1. Take readings of pressure and volume for a suitable range of volumes, determined after a pre-test. Plot a graph as you go.
2. Look for a pattern in the results and then plot a presentation graph, showing the pattern clearly.
3. Are there any regions of the graph that do not fit your pattern as well? Can you account for these deviations in ways that relate back to the state of the gas at that point – or other likely weaknesses in the experimental arrangement?

You have

1. Measured how the volume of a gas changes with its pressure.
2. Thought carefully about why the experiment is set up in a particular way.
3. Produced a set of presentation-quality graphs describing the relationships you found.

Practical advice

This is a straightforward and well known experiment – but it is important that students think through the reasons why the experiment came to be in this form – and consider what could be done to modernise it.

External reference

This activity is taken from Advancing Physics chapter 13, 20E

TAP 601- 5: Changes in temperature, changes in pressure

Common behaviour

Gases are remarkable because they are all so similar. Solids vary considerably because they are tightly bound, and the details of the bonding between the particles affect the properties of the material. Gases are not tightly bound, so are much simpler. Increasing either the mass or speed (or both) of the particles that make up the gas is likely to result in more violent collisions, so increasing the pressure of the gas. It is best to do one thing at a time, so as to keep the story simple, so keep the mass of the particles, and the number of particles, constant.

Increasing the temperature increases the speed of the particles, so expect this change to increase the pressure. Here you are asked to find as exact a relationship as you can between temperature and pressure.

You will need

- ✓ beaker, 400 cm³
- ✓ pressure sensor
- ✓ thermometer
- ✓ gas sample

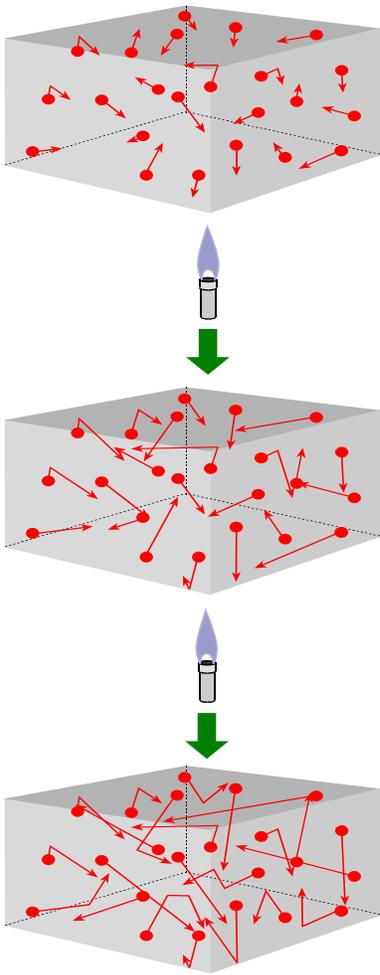


Safety

Strong heating of glass is always hazardous, so protect your eyes, and, in the second example, use a safety screen.

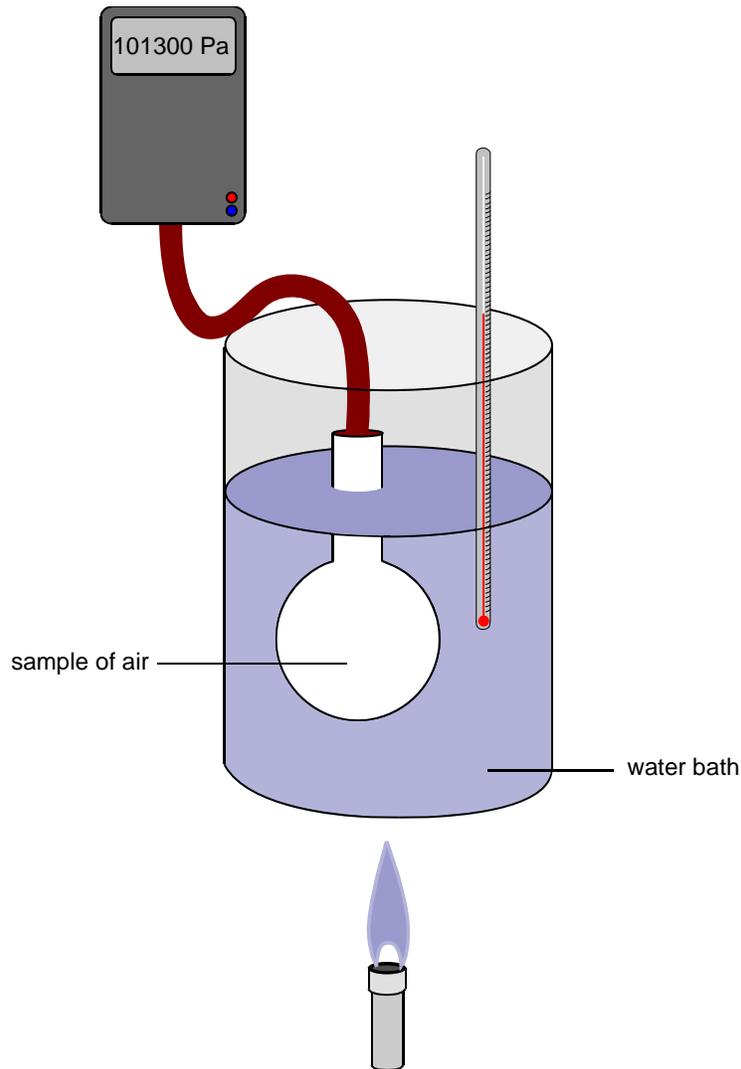
Thinking about the measurements

You need to measure how the temperature of a fixed number of particles affects the pressure. In a school laboratory you may be constrained by the range of temperatures you can reach (remembering that all of the fixed mass of gas must be at one temperature), or the resolution of the pressure meter over this range of temperatures.



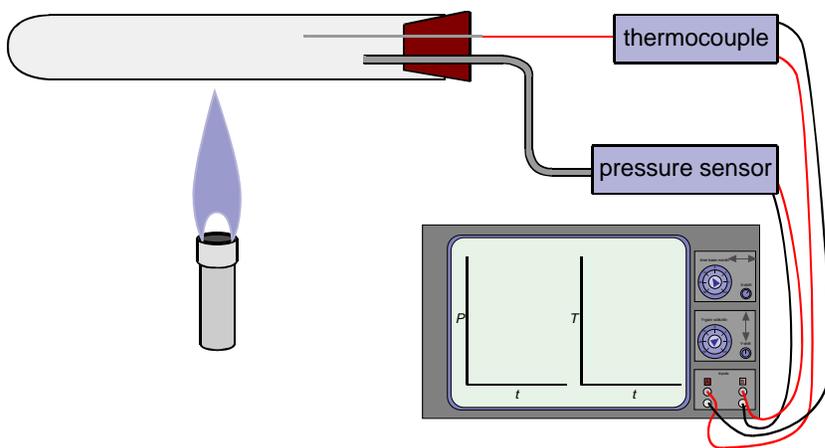
Two solutions to this problem are given below.

A traditional solution



1. Take readings of pressure for the widest range of temperatures that you can manage (0–100 °C?). This is only likely to give a small range of pressures (think about the range of temperatures the gas could exist at). Plot a graph as you go.
2. Look for a pattern in the results and then plot a presentation graph, showing the pattern clearly.
3. Are there any regions of the graph that do not fit your pattern as well? Can you account for these deviations in ways that relate back to the state of the gas at that point – or other likely weaknesses in the experimental arrangement?

A more sophisticated solution



1. Strongly warm the air in a Bunsen flame. Remove the boiling tube from the flame and start capturing data, perhaps for 10 minutes.
2. Once you have checked that your plots of pressure / time and temperature / time are sensible, try plotting pressure / temperature. Look for a clear relationship.
3. Any of your plots may have revealed regions where the readings are not sensible, and you may be able to relate this to the particulars of the experiment that you have just carried out. Once you have eliminated these, produce some presentation-quality graphs.

You have

1. Measured how the pressure of a gas changes with its temperature.
2. Thought carefully about why the experiment is set up in a particular way.
3. Produced a set of presentation-quality graphs describing the relationships you found.

Safety

The most likely failure is the bung being blown out. The second most likely is the boiling tube breaking. A safety screen or screens should be placed to protect the experimenter from both modes of failure.

Practical advice

In this activity it is important that students do not simply collect data and then plot a graph. They also need to be encouraged to appreciate a little of the ingenuity that has resulted in the apparatus being put together in the way that it has. It is also important that students imagine the gas molecules at work, increasing the pressure by blindly colliding with the container walls more and more often as the temperature rises. Producing a datalogged version of this experiment is not so hard. You need to ensure that the seal on the boiling tube is good, as gaining or losing molecules will be fatal. The tube will get hot, so hold it near the bung. Another possible source of error is that not the entire sample of gas is at the same temperature – so place the temperature sensor carefully.

External reference

This activity is taken from Advancing Physics chapter 13, 30E

TAP 601- 6: Changes in temperature, changes in volume

Common behaviour

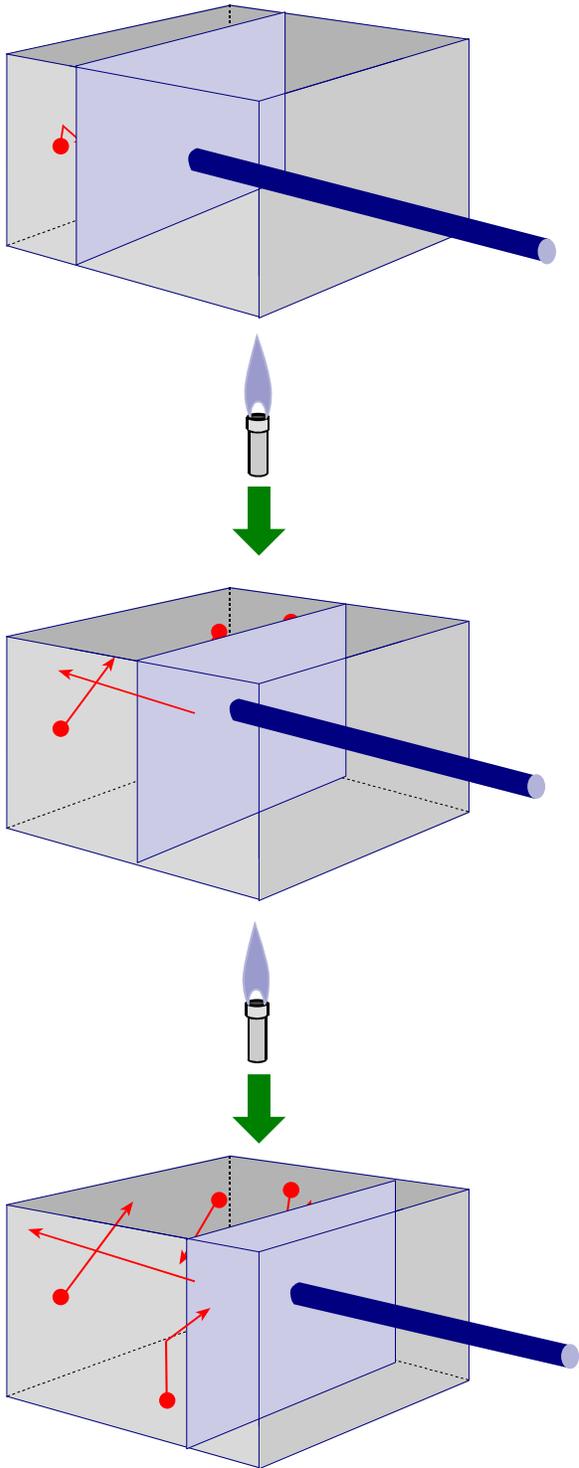
Gases are remarkable because they are all so similar. One of their similarities is that they occupy all the volume that they are placed in, exerting pressure on all of the walls. If the container is flexible then warming the gas can change the volume of the container. Pressure increases as the gas is warmed, then decreases as the molecules become more spread out when the container expands. The walls of the container come to equilibrium again when the pressure inside the container is equal to the atmospheric pressure on the outside of the walls.

Increasing the temperature increases the speed of the particles, so expect this change to increase the volume. Here you are asked to find as exact a relationship as you can between temperature and volume.

You will need

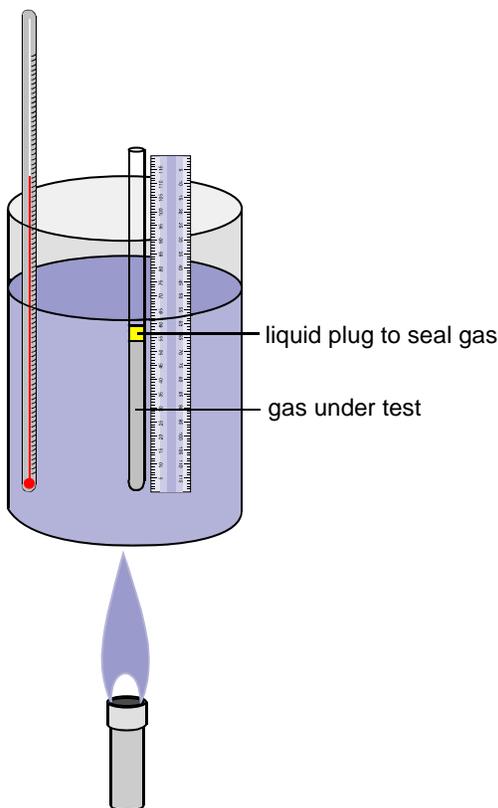
- ✓ Charles' law apparatus

Thinking about the measurements



You need to measure how the temperature of a fixed number of particles affects the volume. In a school laboratory you may be constrained by the range of temperatures you can reach (remembering that all of the fixed mass of gas must be at one temperature), and by the need to have a container for the gas which allows the volume to change, but which does not leak gas. The traditional solution uses a liquid to trap the gas and to provide a flexible wall to the container so allowing changes in volume to be measured.

A traditional solution



1. Take readings of volume for the widest range of temperatures that you can manage (0–100 °C?). This is only likely to give a small change in volume. Plot a graph as you go.
2. Look for a pattern in the results and then plot a presentation graph, showing the pattern clearly.
3. Are there any regions of the graph that do not fit your pattern so well? Can you account for these deviations in ways that relate back to the state of the gas at that point – or other possible weaknesses in the experimental arrangement?
- 4.

Working towards a neater solution

1. You need, if possible, to find a wider range of temperatures.
2. Aim to record the volume automatically, perhaps using a potential divider arrangement connected to a syringe.

3. Check your apparatus for leaks.
4. You will probably find that data are gathered over time – so try plotting temperature / time and volume / time to make sure that the data gathered make sense before processing the information.
5. Any of your plots may have revealed regions where the readings are not sensible, and you may be able to relate this to the particulars of the experiment that you have just carried out. Once you have eliminated these, produce some presentation-quality graphs.

You have

1. Measured how the volume of a gas changes with its temperature.
2. Thought carefully about why the experiment is set up in a particular way.
3. Produced a set of presentation-quality graphs describing the relationships you found.

Safety

The traditional apparatus uses a narrow bore tube but not a capillary tube, e.g. a bore of 1 to 1.5 mm. It is now difficult to obtain tubing of this size and a bore of 2 to 3 mm is used.

The air was traditionally trapped by a short thread of concentrated sulphuric acid to ensure that the air is dry. The acid drains down a tube with wider bore than that described above. Consequently, the experimenter is tempted to refill the tube just before the experiment.

In this case, the tube must be heated in an oven to about 200 °C, held upside down with its open end under the surface of the acid in a small test tube or beaker until it has cooled sufficiently to draw a small thread from the acid and allowed to return to room temperature.

A face shield or goggles must be worn during this procedure.

Practical advice

In this activity it is important that students do not simply collect data and then plot a graph. They also need to be encouraged to appreciate a little of the ingenuity that has resulted in the apparatus being put together in the way that it has. It is also important that students imagine the gas molecules at work, producing the increase in volume by blindly colliding with the container walls more and more often as the temperature rises.

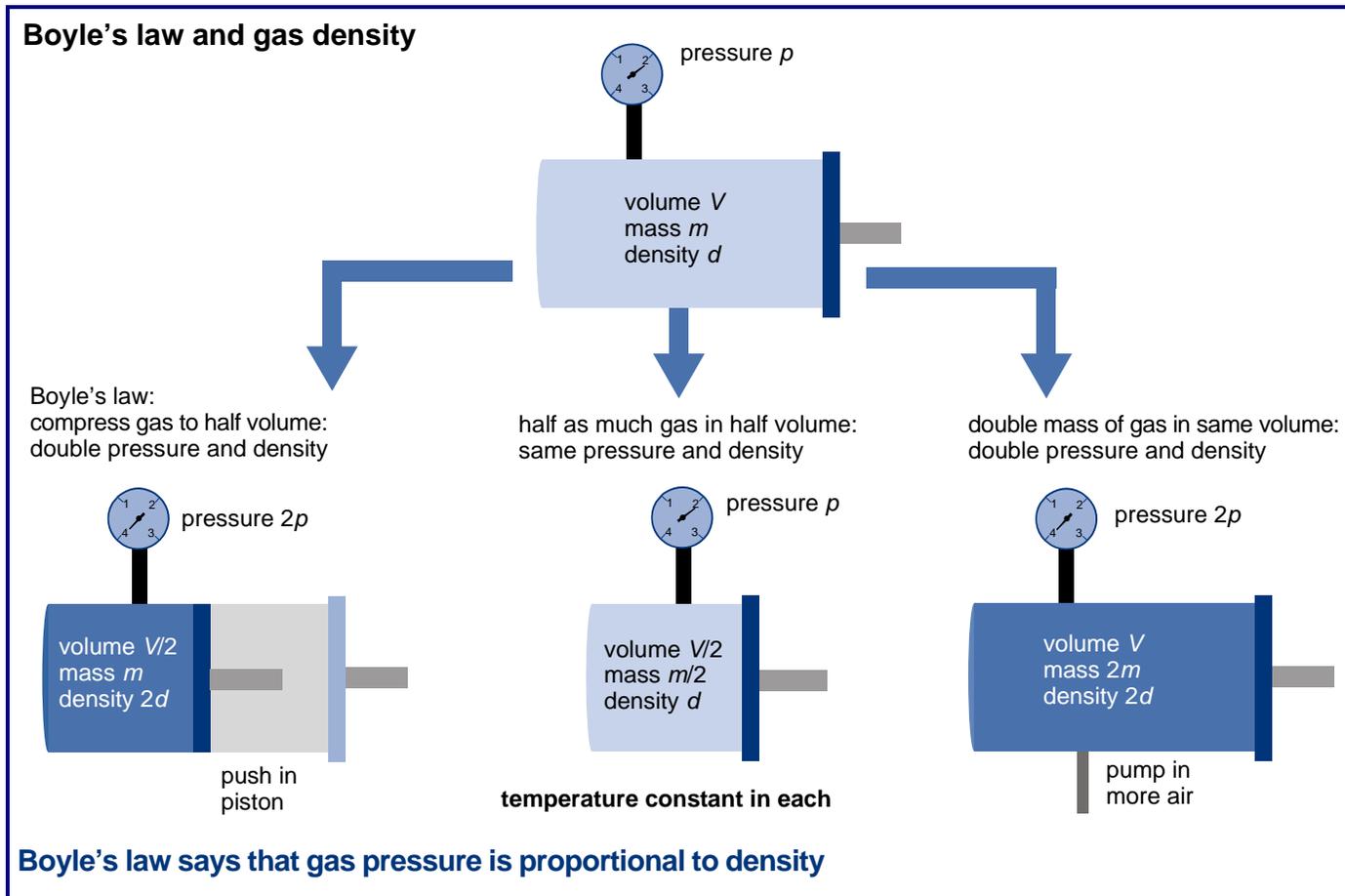
Alternative approaches

Producing a datalogged version of this experiment is not easy. The problems revolve around the measurement of volume, where a linear position sensor allied to a syringe look attractive. However, the difficulties in making the syringe airtight result in considerable discrepancies.

External reference

This activity is taken from Advancing Physics chapter 13, 40E

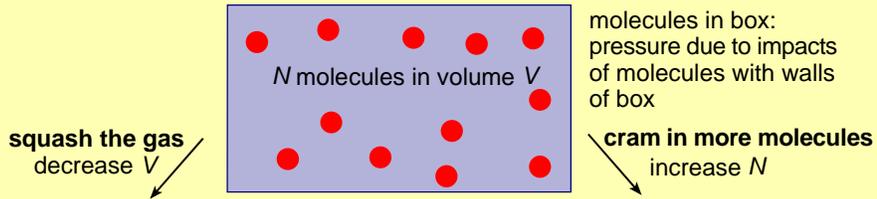
TAP 601-7: Boyle's law, density and number of molecules



This diagram shows how pressure and density are connected, providing a basis for understanding where Boyle's law comes from.

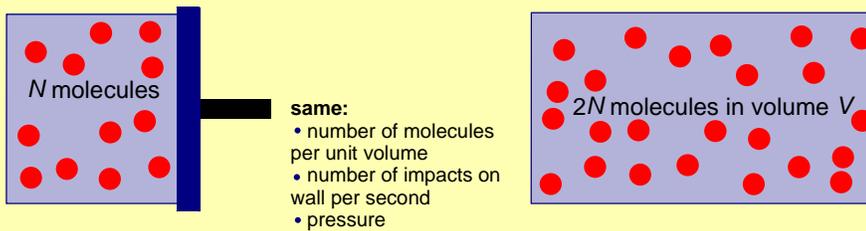
Boyle's law and number of molecules

Two ways to double gas pressure



piston squashes up same molecules into half the volume, so doubles the number per unit volume

add extra molecules to double the number, so double the number per unit volume



pressure proportional to $1/\text{volume}$
 $p \propto 1/V$

pressure proportional to number of molecules
 $p \propto N$

If... pressure is proportional to number of impacts on wall per second

and if... number of impacts on wall per second is proportional to number of molecules per unit volume

Then... pressure is proportional to number of molecules per unit volume

$$p = \text{constant} \times N/V$$

Boyle's law in two forms

$$pV = \text{constant} \times N \quad p = \text{constant} \times N/V$$

Boyle's law says that pressure is proportional to crowding of molecules

This diagram shows how pressure and number of molecules are connected, providing a basis for understanding where Boyle's law comes from.

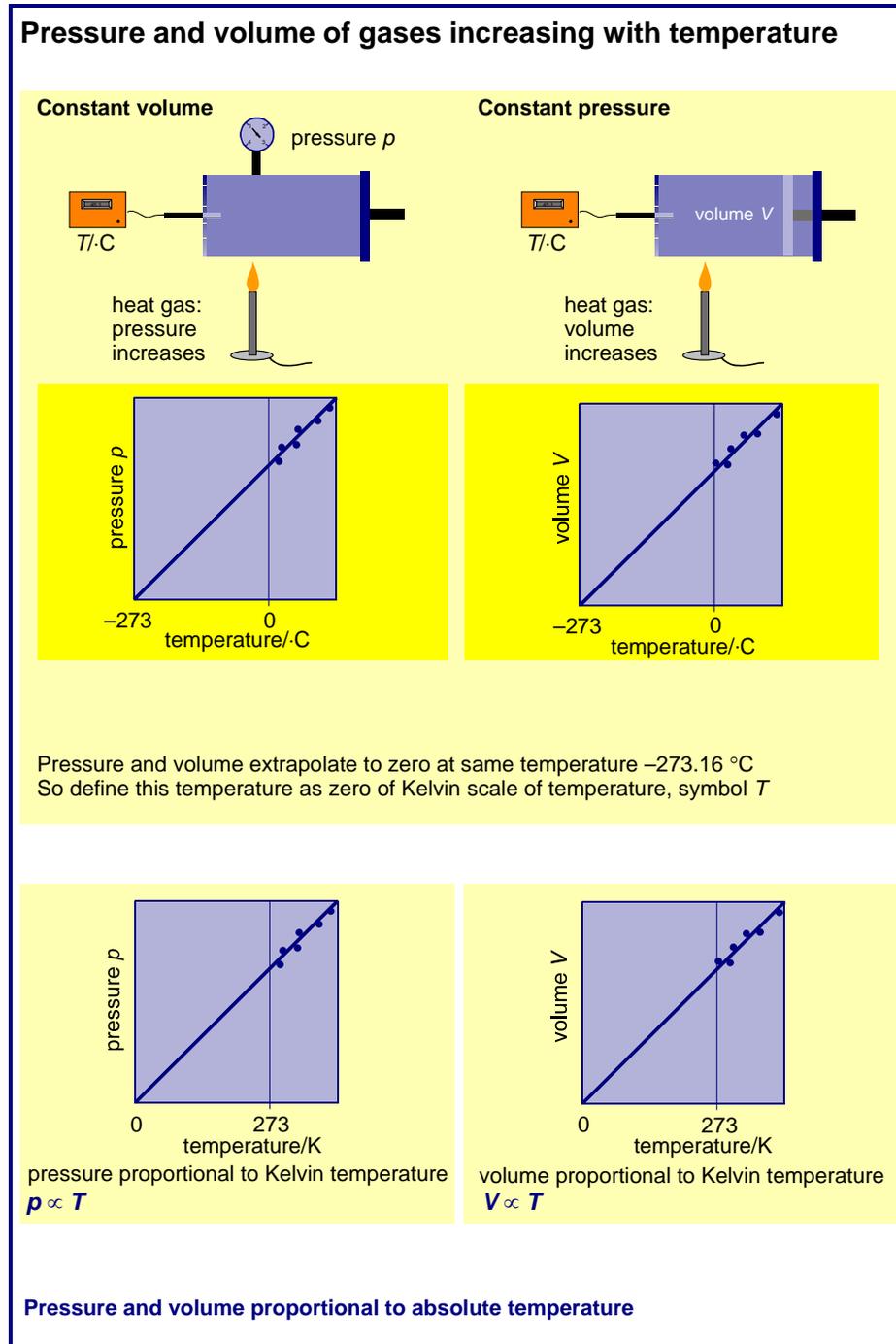
Practical advice

These diagrams are reproduced here so that you can discuss them with your class.

External reference

This activity is taken from Advancing Physics chapter 13, 100

TAP 601- 8: Changing pressure and volume by changing temperature



This diagram shows how temperature affects pressure and volume, suggesting extrapolation to absolute zero. In turn this provides a basis for the absolute zero of temperature.

Practical advice

These diagrams are reproduced here so that you can discuss them with your class.

External reference

This activity is taken from Advancing Physics chapter 13, 200