SCOPE, SEQUENCE, and COORDINATION

A National Curriculum Project for High School Science Education

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Student Materials

Learning Sequence Item:

929

Gas Laws

March 1996
Adapted by: Dorothy Gabel

Contents

This micro-unit was adapted by **Dorothy Gabel** (Indiana University, Bloomington)

Science as Inquiry

How Much Space?

How much space is there between gas molecules?

From the measurements of the piece of solid dry ice, calculate the volume of the piece of dry ice. After the dry ice has been dropped into the warm water, calculate the volume of the gas produced. Now compare the volume of the gas with that of the solid.

1. What is the ratio of the carbon dioxide gas to that of the carbon dioxide as a solid (dry ice)?

2. Explain why there are differences between the volumes of the solid and the gas.

3. Using the data collected, calculate the relative distances between the molecules of carbon dioxide in the gaseous state.

4. Using dots to represent molecules of carbon dioxide, draw pictures to represent the same volume of carbon dioxide as a solid and as a gas.

5. Compare the compressibility of solids with that of gases and explain differences.

6. Predict the compressibility of liquids as compared to that of solids and gases. Explain in terms of the particle nature of matter.

Science as Inquiry

Eating Space?

What happens to marshmallows when pressure is reduced?

A marshmallow is placed on the platform of a vacuum pump under a glass dome. The pump is turned on.

1. Predict what you think will happen when the pump is turned on.

2. Describe your observations after the pump is turned on.

3. Explain your observations of what is happening to the air in the dome and to the marshmallow in terms of the particulate nature of matter.

4. What are the air molecules doing to produce the effect in the marshmallow?

5. If you ate the marshmallow after the air was evacuated would you consume more calories than before? Explain your answer.

6. From your observations, explain the relationship between the volume and pressure of gases.

Science as Inquiry

Reduce the Pressure

What happens to balloons when pressure is reduced?

An inflated balloon is placed on the platform of a vacuum pump under a glass dome. The pump is turned on.

1. Predict what you think will happen when the pump is turned on.

2. Describe your observations after the pump is turned on.

3. Explain your observations of what is happening to the air in the dome and to the balloon in terms of the particulate nature of matter.

4. Do you think that the inflated balloon would have the same mass before and after the air was evacuated by the pump?

5. From your observations, explain the relationship between the volume and pressure of gases.

Science as Inquiry

Inflating the Tube

Why does the tube inflate?

Observe the toothpaste tube before, during and after heating.

1. Describe your observations before, during, and after the toothpaste tube was heated.
2. Did the mass of the air in the tube change during your observations? Explain.
3. Explain what happened in terms of the air particles inside the tube.
4. Did the density of the air change as it was heated? Explain.
5. From your observations, explain the relationship between the volume and temperature of gases.

Science as Inquiry

Cartesian Diver

What makes the eyedropper sink?

Observe the Cartesian diver. Press on the sides of the bottle and observe. Release the pressure and observe.

1. Describe your observations before you made the diver sink, after it sank, and after it returned to the top.
2. Did the mass of the air in the diver change during your observations? Explain.
3. Explain how the Cartesian diver works.
4. Did the density of the air change when the diver sank? Did the density of the water change? Explain.

Science as Inquiry

Balloon Volumes

How is gas volume related to temperature?

Observe the volume of the balloons under the various conditions.

1. Describe your observations when the balloon is placed in the hot water and in the ice water.
2. Did the mass of the air in the balloon change during your observations? Explain.
3. Did the density of the air change during your observations?
4. Explain what was happening to the molecules of gas during the activity?
5. What was happening to the pressure of the gas inside the balloons during the activity? Explain.

Student Sheet

Science as Inquiry

Filling Bottles

How is gas volume related to pressure?

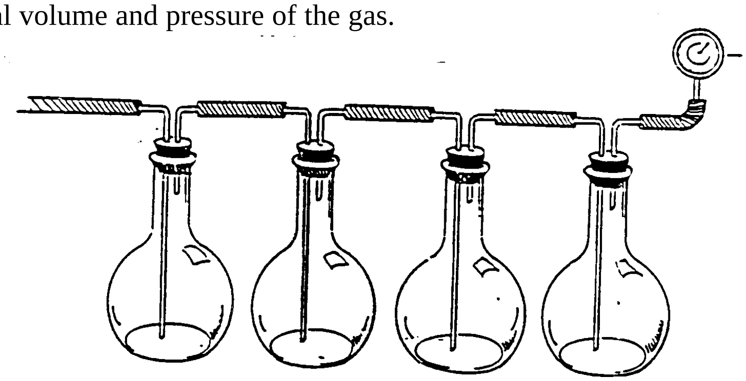
Overview:

Four bottles are connected to a pressure gauge. Water is poured into the first bottle at definite intervals and pressure readings are made. Is there a relationship between the pressure and the volume of a gas?

Procedure:

Before assembling the apparatus like the one shown below, determine the volume of one of the flasks. You will need to know the volume at 1/4, 1/2, and 3/4 full as well as the volume of the full flask. Determine these volumes, record them in your data table, and mark the flasks with a marking pencil or tape.

Assemble the flasks as shown in the diagram. Secure the flasks with ring stands and clamps. Wrap wire around the stoppers and the necks of the bottles to secure the stoppers as champagne corks are secured to their bottles. Attach one end of the tubing to a pressure gauge and the other end to the water faucet. Have your instructor check your setup before beginning the experiment. Make a data table and record the initial volume and pressure of the gas.



Now slowly add water to the first flask until it is 1/4 full. Record the volume and temperature. Repeat at the 1/2, 3/4 and full intervals.

1. Examine your data. What qualitative statement can you make about the relationship between the pressure and volume of a gas?

2. From your data determine if there is a mathematical expression that can be written that expresses the relationship between the pressure and the volume of the gas. (Hint: Try adding, subtracting, multiplying, or dividing the variables to determine if a constant can be obtained.)

3. Graph the pressure versus volume. Does your graph indicate that a relationship exists between pressure and volume?

4. Account for some of the experimental errors that might occur in the collection of data.

Student Sheet

Science as Inquiry

Pressure-Volume

Is there a mathematical relationship between the pressure and the volume of a gas when the temperature remains constant?

Overview:

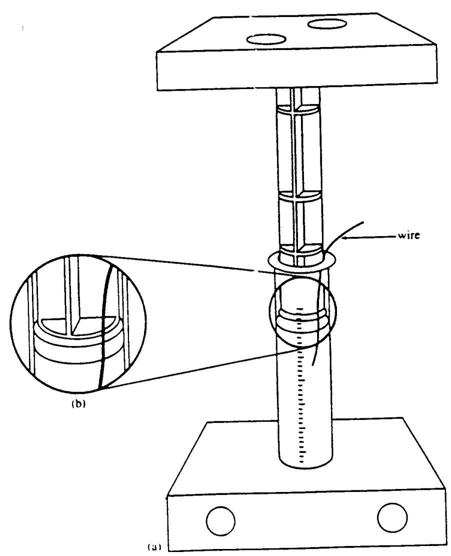
Air is placed in a syringe supported on a wooden base. Masses are placed on a wooden block on top of the syringe plunger and differences in volume are noted.

Procedure:

Set up the apparatus as shown in the diagram.

First adjust the volume of gas to read 30 cm^3 by placing a string next to the plunger and adjusting the plunger until it reads 30 cm^3. Then remove the string.

Make a data table to record the pressure (in bricks) and the volume. Place one brick on the system and record the new pressure and volume. Repeat for four or five bricks. Because there is some friction involved in the system, you may wish to also measure the volume and pressure as you remove the bricks. Record the data.



1. How might you reconcile differences in the pressure readings when adding and removing bricks?

2. Graph the data. Do the data indicate that the pressure is inversely proportional to the volume? How can you tell?

3. Use your data to determine if a mathematical operation (addition, subtraction, multiplication, or division of the two variables) results in a constant value. Show the relationships you tried in table form. There is no need to show the arithmetic because you undoubtedly will use a calculator. Does a relationship exist? Which relationship is most probable?

4. What might account for experimental errors in this experiment?

5. In your calculations in question 4, you used bricks as pressure. Is additional pressure from other sources being exerted on the syringe?

6. Adjust your graph to represent an inverse relationship. Why must it be adjusted?

7. If a 100-mL sample of gas at 25 °C exerts a pressure of 700 torr, what volume would the gas occupy at the same temperature if the pressure was reduced to 600 torr?

Science as Inquiry

Temperature-Volume (1)

Is there a mathematical relationship between the temperature and the volume of a gas when the pressure remains constant?

Overview:

Air is placed in a pipette and its volume is measured by determining the number of drops of water the pipette contains. Pipettes of air are heated by placing them in water baths at various temperatures. The volume of the gas at elevated temperatures is measured, and the relationship between volume and temperature at constant pressure is obtained.

Procedure:

Measure the volume of a pipette by completely filling it with water and counting the drops that are released from the pipette until it is empty. Assume that all the pipettes you will use are of equal volume. Record the volume in drops and the room temperature.

In this activity you will record the volume and the temperature of air at constant pressure and various temperatures. You will need two beakers of water. One you will maintain at room temperature and the other you will heat as directed by your teacher. You will use it as a water bath into which you place your pipette containing air so that you will be able to measure air volumes at several elevated temperatures. Do not heat the water over 90 °C to avoid melting the pipette. You will need to use dry pipettes for each reading that you take. Why?

As the air is heated it will expand and some will leave the pipette. You can determine how much has left by bending the tip of the pipette while it is still in the warm water bath, and then submerging the entire pipette in the room temperature bath. When you release the tip, water will go into the pipette. Why? If you measure the volume of water that goes into the pipette and add this to the original volume of air at room temperature, this new volume represents the volume of the gas at the elevated temperature. Measure the volume of air at several different temperatures.

1. Make a graph of your findings. Place the volume of air on the vertical axis and the temperature on the horizontal axis. Extrapolate the temperature so it crosses the x-axis.

2. At what temperature does the gas have no volume? In reality, the gas at this temperature would have changed to a solid. This temperature is called absolute zero. Theoretically, it can never be reached.

3. Adjust your scale on the graph by marking the absolute zero as zero and converting the other Celsius temperatures to the Kelvin scale. This will make all of your temperatures positive. The abbreviation for Kelvins is K (note that no degree sign used).

4. Using your graph, write a mathematical equation expressing the relationship between volume (V) and temperature (T) when the pressure is constant.

5. Express qualitatively the relationship between the volume and temperature of a gas when the pressure is constant.

6. Find the volume that 600 mL of hydrogen gas would occupy if its temperature were changed from 30 °C to 90 °C and the pressure remained constant.

7. To what temperature would you need to heat 500 liters of oxygen gas at room temperature (25 °C) to change its volume to 2000 liters?

Student Sheet

Science as Inquiry

Temperature-Pressure

Is there a mathematical relationship between the temperature and the pressure of a gas when the volume remains constant?

Overview:

Air is placed in a flask of definite volume. The flask is submerged in water at various temperatures and the pressure exerted by the air in the flask is measured on a gauge. The relationship between pressure and temperature at constant volume is obtained.

Procedure:

Secure the pressure gauge to the flask by wrapping the stopper with wire as demonstrated by your teacher. Take an initial reading of the temperature and the pressure of the air in the flask. Then half fill a large container with a mixture of ice and water, and another with boiling water. Submerge the flask in each of the containers for about five minutes until you see no further change on the pressure gauge. Record the data.

1. Make a graph of your findings. Place the pressure of air on the vertical axis and the temperature on the horizontal axis. Extrapolate the temperature so it crosses the x-axis.

2. At what temperature does the gas exert no pressure? In reality, the gas at this temperature would have changed to a solid. This temperature is called absolute zero. Theoretically, it can never be reached.

3. Adjust your scale on the graph by marking the absolute zero as zero and converting the other Celsius temperatures to the Kelvin scale. This will make all of your temperatures positive. The abbreviation for Kelvins is K (note that no degree sign is used). The abbreviation for temperature in Kelvins is T.

4. Using your graph, write a mathematical equation expressing the relationship between pressure (P) and temperature (T) when the volume is constant.

5. Express qualitatively the relationship between the pressure and temperature of a gas when the volume is constant.

6. Find the pressure that 600 mL of standard pressure hydrogen gas would exert if its temperature were changed from 30 °C to 90 °C and the volume remained constant.

7. To what temperature would you need to heat 500 liters of oxygen gas at STP to change its pressure to 1000 torr?

Student Sheet

Science as Inquiry

Temperature-Volume (2)

Is there a mathematical relationship between the temperature and the volume of a gas when the pressure remains constant?

Overview:

The volume of air is trapped inside a capillary or small-diameter glass tube. It is attached to a thermometer and placed in water at various temperatures. As the temperature of the water bath increases, the trapped air expands. By measuring the diameter of the tube and the height of the column of air, the volume can be calculated over a temperature range of 100 degrees.

Procedure:

In this activity you will record the volume and the temperature of air that has been trapped in a capillary tube at constant pressure and various temperatures.

As the air is heated in the capillary tube it will expand and its height can be measured. This height can be converted to a volume by multiplying the cross-sectional area by the height. Volume = height in cm × cross-section area (p$r^{2)}$ in cm^2.

Prepare a data table to record the height and the volume of the column of air at various temperatures. First measure the height of the air column when the tube is placed in ice water, and then take other readings as the tube is heated from room temperature to that of boiling water.

1. Make a graph of your findings. Place the volume of air on the vertical axis and the temperature on the horizontal axis. Extrapolate the temperature so it crosses the x-axis.

2. At what temperature does the gas have no volume? In reality, the gas at this temperature would have changed to a solid. This temperature is called absolute zero. Theoretically, it can never be reached.

3. Adjust your scale on the graph by marking the absolute zero as zero and converting the other Celsius temperatures to the Kelvin scale. This will make all of your temperatures positive. The abbreviation for Kelvins is K (note that no degree sign is used).

4. Using your graph, write a mathematical equation expressing the relationship between volume (V) and temperature (T) when the pressure is constant.

5. Express qualitatively the relationship between the volume and temperature of a gas when the pressure is constant.

6. Find the volume that 600 mL of hydrogen gas would occupy if its temperature were changed from 30 °C to 90 °C and the pressure remained constant.

7. To what temperature would you need to heat 500 liters of oxygen gas at room temperature (25 °C) to change its volume to 2000 liters?

Science and Technology

Gas Laws and Scuba Diving

What happens if scuba divers hold their breath while making emergency ascents to the surface from depths of 30 meters or more?

Why shouldn't divers fly or take hot showers soon after deep dives?

Is contaminated compressed air more dangerous to the diver at the surface or at a depth of 30 meters?

Pressure

We live in a sea of air. Since air molecules constantly bombard us, we always experience a pressure of about 760 mm of mercury (or one atmosphere) at the Earth's surface. This is equivalent to 14.7 lb on each square inch of surface. If we zoom to the top of a tall building in an elevator we are no longer as deep in the sea of air as at ground level and, therefore, the pressure around us becomes lower. Ears are usually the first to respond to this change. Wiggling your jaw or swallowing sometimes corrects any discomfort or strange sensations in the ear by opening the tubes connecting the inner ear and throat, allowing the inside pressure to equalize with the outside. A reverse pressure effect is obvious during a rapid airplane descent or during a drive from a mountain pass to the valley floor below.

Divers are surrounded by water molecules in constant motion that exert pressure on their bodies. When you dive to the bottom of the deep end of a swimming pool, you feel a great deal of pressure exerted by the water. Because water is much more dense than air, pressure changes are much greater for a given change in depth in water than for the same depth change in air. For example, water exerts over 100 lb of force on the surface of a one gallon metal can pushed just one foot below the water surface. If the metal can contains air, it would not have to be pushed very far below the water surface before the can would start to collapse due to water pressure. Can divers be crushed by the pressure of water in the same manner as the can if they go too deep? After all, for every 10 meters (about 33 ft) in depth, divers experience an additional pressure of one atmosphere.

Pressure-Volume Effects

The changes in pressure experienced by divers are most noticeable on body cavities that contain air, such as the lungs, the middle ear, and the sinus cavities. Boyle's law describes how these gas volumes respond to changes in pressure. For a constant does amount of gas at a constant temperature, Boyle's law states: The volume of a gas sample varies inversely with its pressure.

If divers descend without scuba gear, the amount of gas contained in their body cavities is constant and the volume of these cavities decreases as the surrounding water pressure becomes greater. However, this crushing effect or squeeze is not experienced by divers using scuba gear because the regulator on their air tanks delivers air at the same pressure as the surroundings.

This means that the air in divers' lungs is at a pressure equivalent to four atmospheres at a depth of 30 meters. If divers must make emergency ascents from this depth they must remember to breathe out regularly as they return to the surface. If they don't, the pressure of the air in their lungs will cause their lungs to expand. The extreme distortion of the lungs can cause some of the alveoli (the small sacks in the lungs) to rupture. If this happens, air can enter the bloodstream and cause a blockage that may lead to a variety of problems including loss of consciousness, brain damage, and heart attacks.

The rate of lung expansion increases dramatically as the divers ascend. According to Boyle's law the volume of a flexible gas container will approximately double when the surrounding pressure decreases to one-half its original value. If the divers ascend while holding their breath from a depth of 30 meters (where the pressure is about four atmospheres), their lungs would have to double in volume when they are at 10 meters (where the pressure is about two atmospheres) to equalize the pressure of the water. Of course, this does not happen because the lungs are contained by the rib cage and the muscle system, and the divers are forced to breathe out.

Pressure-Solubility Effects

Not only does the pressure affect the volume of trapped gases, it also influences the solubility of gases in liquids. Divers must be aware of the principles described by Henry's law, which states: The amount of gas that will dissolve in a liquid at a given temperature varies directly with the pressure above the liquid.

Henry's law is useful, therefore, in explaining why during a dive any gases entering the lungs are absorbed to a greater extent in the diver's blood. Although this increased solubility of gases in the blood may create no problems during the dive, the diver's body experiences an effect similar to opening a can of soda when the diver ascends rapidly to the surface. This effect can be accentuated if the diver takes a high-altitude plane flight soon after a dive. In particular, nitrogen gas bubbles that form in the blood and other body fluids can produce a multitude of problems.

These problems depend on the location of the gas bubbles, the size and number formed, and the way they are transported by the diver's circulatory system. The bubbles can cause localized pain, itching of the skin, breathing difficulty, and can lead to paralysis, unconsciousness, and death.

To minimize gas bubble formation (decompression sickness or "the bends"), divers carefully follow tables prepared by the U.S. Navy that describe the time limits for dives at various depths greater than 10 meters. The essence of the process described by the tables involves ascending to a certain point and then remaining at that depth for a time period to allow some of the dissolved nitrogen to escape.

Depending on the initial depth, there may be several of these "hold points" during the ascent. If divers experience decompression sickness, the only mode of treatment is to put them in a decompression chamber, increase the pressure surrounding their bodies, and slowly decompress them back to one atmosphere of pressure.

The increased solubility of nitrogen gas at higher pressures may also have a narcotic effect. Nitrogen narcosis or "rapture of the deep" generally does not occur until divers reach depths of about 30 meters. The symptoms are similar in nature to intoxication by alcohol. The divers have a feeling of happiness, overconfidence, tingling or numbness in their arms or legs, and memory impairment. This narcotic effect of nitrogen is just divers should never work alone underwater.

Another application of Henry's law involves contaminants such as carbon monoxide (CO) that might be present in the compressed air used by divers. Of course, every attempt is made to ensure the purity of the air in scuba tanks, but if a contaminant is present to the extent of just 1%, its presence is more serious during a dive. For example, at a depth of 40 meters, the pressure is equivalent to about five atmospheres.

Because the regulator divers air at the same pressure as the surroundings, each breath contains five times more contaminant molecules than each breath from that same tank at the surface. This is equivalent to breathing air containing 5% of that contaminant at the surface.

As the pressure increases during a dive, the solubility of oxygen in the blood also increases proportionately. This means that the effects of poisoning by a trace of carbon monoxide contaminant may go unnoticed during a dive since sufficient oxygen is available for normal cellular respiration. However, as divers surface, the solubility of oxygen decreases in their bloodstreams.

Because the carbon monoxide-hemoglobin combination is so stable, there may not be a corresponding decrease of carbon monoxide in the blood. If the divers do not have enough hemoglobin available to bond with oxygen cell respiration, they may lapse into unconsciousness.

Temperature-Solubility Effects

Gas solubility is also affected by changes in temperature. Have you ever noticed that as a cold glass of water warms to room temperature, air bubbles form, clinging to the inside of the glass surface? These bubbles are composed of air that was dissolved in the cooler water. Can you use this information to explain why it is dangerous for a diver to take a hot shower after a deep dive?

A scuba diver with a good basic understanding of gas behavior will better appreciate what is happening during a dive. If you are a scuba diver, this understanding could save your life!

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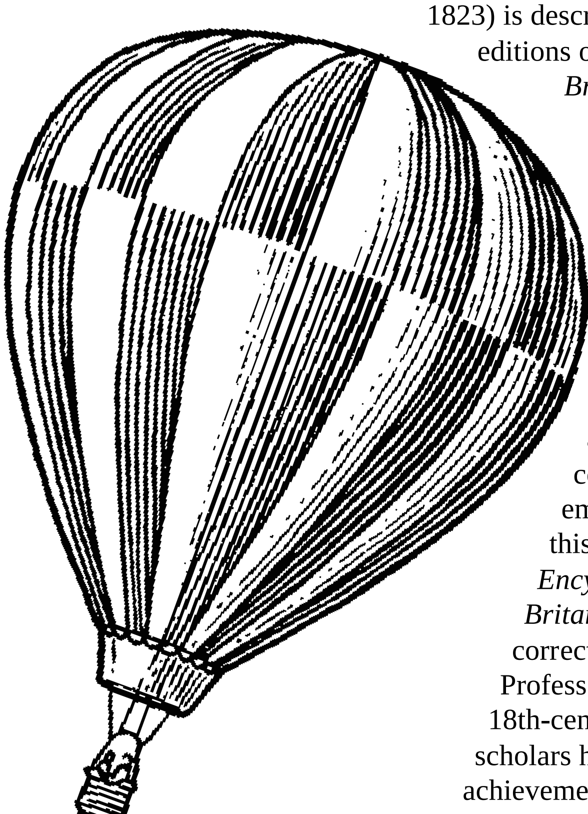
History and Nature of Science

How the Right Professor Charles Went Up in the Wrong Kind of Balloon



The chemical equivalent of "Who's buried in Grant's Tomb?" might well be "Who discovered Charle's law?" In this case, however, the answer is not so straightforward. Jacques Alexander Cesar Charles (1746-1823) is described in older editions of the *Encyclopedia Britannica* as a "French mathematician and physicist" who "was elected to the [French] Academy of Sciences" in 1785 and whose "published papers are chiefly concerned with mathematical topics." In this instance the *Encyclopedia Britannica* is only partly correct. There were *two* Professors Charles in late 18th-century Paris, and scholars have credited the achievements of Jacques Charles-le-Geometre (Charles the Obscure) to the more familiar Professor J.A.C Charles (1, 2). This Professor Charles, who is near and dear to the hearts of all chemistry students, was not an able mathematician, published little in that area, and was not elected to the Academy of Sciences until 1795.

Our Professor Charles has two claims to fame: He was the first man to go up in a hydrogen balloon, and he appears to have been the first to demonstrate experimentally that the volume of a fixed mass of gas at constant pressure is directly proportional to the (absolute) temperature. This principle, the expansion of a gas which is heated, is now known as Chales's law and provides the lift hot-air balloons. Did Charles *discover* Charles's law? Well yes, but he did not publish his discovery and, on no less an authority than Michael Faraday, if a discovery is not published it might almost as well not have been made. Charles's law was first published by John Dalton in 1801 and by Joseph Louis Gay-Lussac in 1802. Dalton's experimental data were typically somewhat sloppy, but Gay-Lussac's were considerably

By Derek A. Davenport. Reprinted with permission from *Chem Matters,* December 1983.

more precise than the then-unpublished data of Charles. Nonetheless it was the generous Gay-Lussac who insisted that Charles, his fellow countryman and, as we shall see, fellow-balloonist, should have the honor of priority. In doing so, Gay-Lussac was to save future generations of students from considerable confusion since *he* was later to discover the law of combining volumes, which now goes by his name, and John Dalton was soon to come up with his law of partial pressures.

On June 4, 1783, the brothers Joseph and Etienne Mongolfier, paper manufactures of Annonay in central France, successfully launched the world's first hot-air balloon (3). It was a huge, unmanned contraption made of layers of linen and paper fastened together with buttons. Suspended beneath was a basket of burning straw. Since such balloons are open at the bottom the pressure inside must be virtually the same as that outside, and the lifting power comes from the Charles's law lowering of the density of the hot trapped air. The news of the Montgolfiers' achievement quickly reached Paris but in the process confusion arose as to the nature of the lifting gas they had used. At that time Antoine Lavoisier, who was later to serve on the Academy of Science's Committee on "Aerostation," was still in the process of establishing the composition of the atmosphere. The only known gas of low density was "inflammable air" or hydrogen and J.A.C Charles was commissioned to build a balloon filled with hydrogen. Shortly afterward Etienne Montgolfier in Paris and immediately began the construction of a large hot-air balloon.

The rivalry between the supporters of hydrogen balloons (Charlieres) and hot-air balloons (Montgolfiers) was intense but not unfriendly. Charles won the first round. Because of the superior lifting powers of hydrogen, his balloon was much smaller than that of Etienne Montgolfier, but the problems and dangers of generating and containing huge amounts of hydrogen were immense. By the evening of Aug. 26, 1783, preparations were completed and the hydrogen-filled balloon (accompanied by flaming torches) was moved by night across Paris to the launching site not far from where the Eiffel Tower now stands. Late the following afternoon the unmanned balloon was released, soon disappeared from sight, and flew a distance of 20 km before crashing to earth near a village of Gonesse where it was immediately attacked with scythes and pitchforks by frightened villagers. Among the many distinguished spectators was Benjamin Franklin who, together with John Jay and John Adams, was to sign the Treaty of Paris ending the American War of Independence one week later. When a cynic asked Franklin what possible use the new invention could have, he quietly answered "Of what use is a newborn baby?"

The next two rounds went to the Montgolfiers. On Sept. 15, 1783 Ethienne Montgolfier's magnificently decorated hot-air balloon took off from the Palace of Versailes in the presence of King Louis XVI and Marie Antoinette. Aboard were the first aerial passengers—a cock, a duck, and a sheep. When the balloon landed, the first to reach the scene was the chemist, Pilatre de Rozier. He ????? to be the first man in space. However, Charles was also working on a larger man-carrying hydrogen balloon and so the race was on: the chemist, Rozier, betting on a "physicist's balloon" and the physicist, Charles, betting on a "chemist's balloon." The chemist won! On Nov. 21, 1783, Pilatre de Rozier accompanied by the Marquis d'Arkabdes made the first manned flight, one of 20-minutes' duration-a giant leap compared to the Wright brothers' first flight at Kitty Hawk. Ten days later, on Dec. 1, 1783 Charles, accompanied by balloon-maker Marie-Noel Robert, successfully flew in a hydrogen balloon. They came down safely 40 km from Paris. There Robert stepped out and, as night fell, the right Professor Charles made man's first solo flight—in the wrong king of balloon!

J.A.C. Charles was never to make another ascent. Pilatre de Rozier made several more, and on June 15, 1785, he became aviation's first casualty when he died attempting he second aerial crossing of the English Channel. He had reasoned that since both hydrogen and hot-air balloons had their separate advantages a combination of the two would be even better. Rozier was accustomed to living dangerously—one of his favorite chemical lecture—demonstrations consisted of flushing his lungs with hydrogen and then speaking in the resulting high-pitched voice (today we tend to use helium). The final flourish (today we would tend to omit this!) was to light the hydrogen as it issued from his mouth *(4)*. Such a man was obviously the "Right Stuff" to fly a hybrid hot-air balloon. Alas, his luck ran out, and he and a companion crashed shortly after takeoff from Boulogne.

And Gay-Lussac the balloonist? In 1804 he went up alone in a hydrogen balloon and without oxygen reached an altitude of 7016 meters, a record that was to last for over 40 years.

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