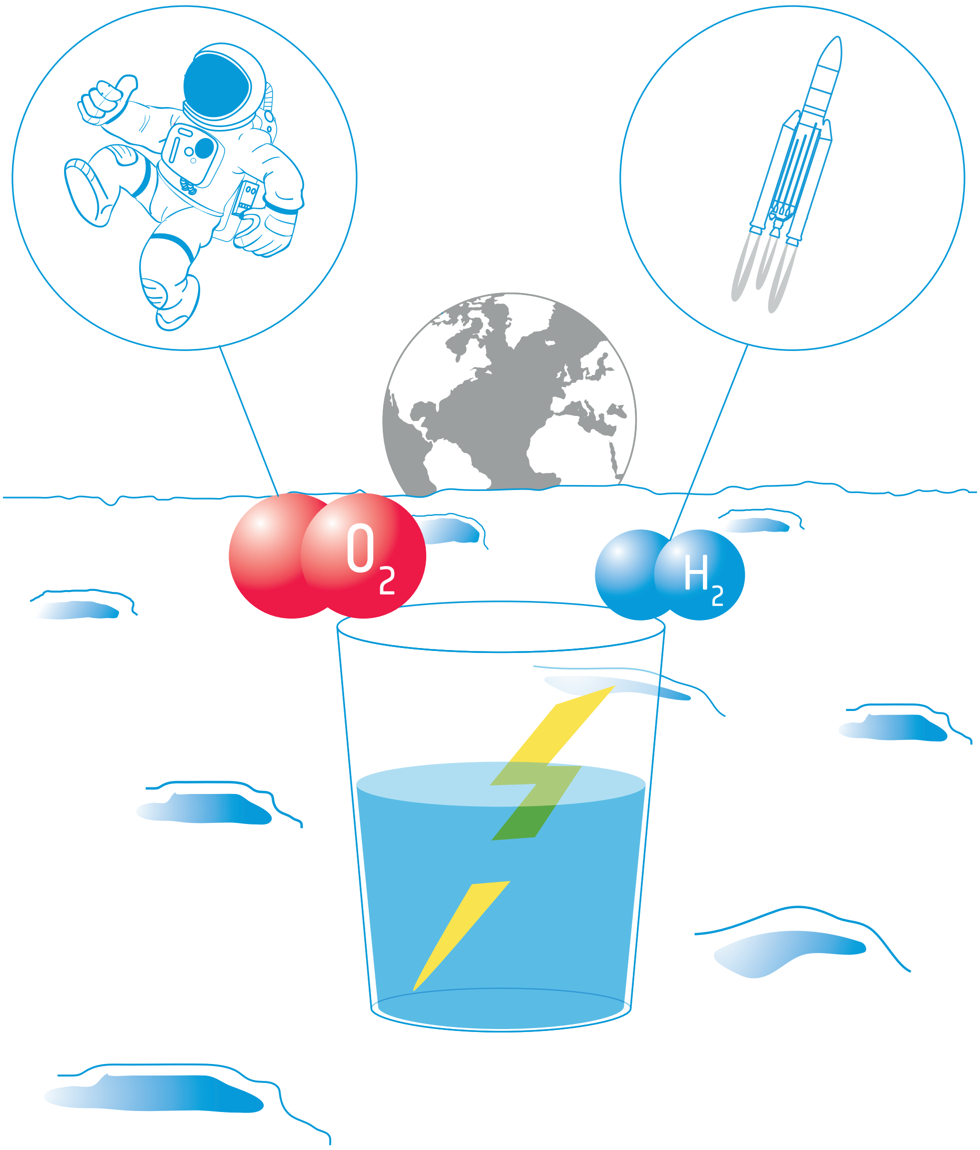
teach with space

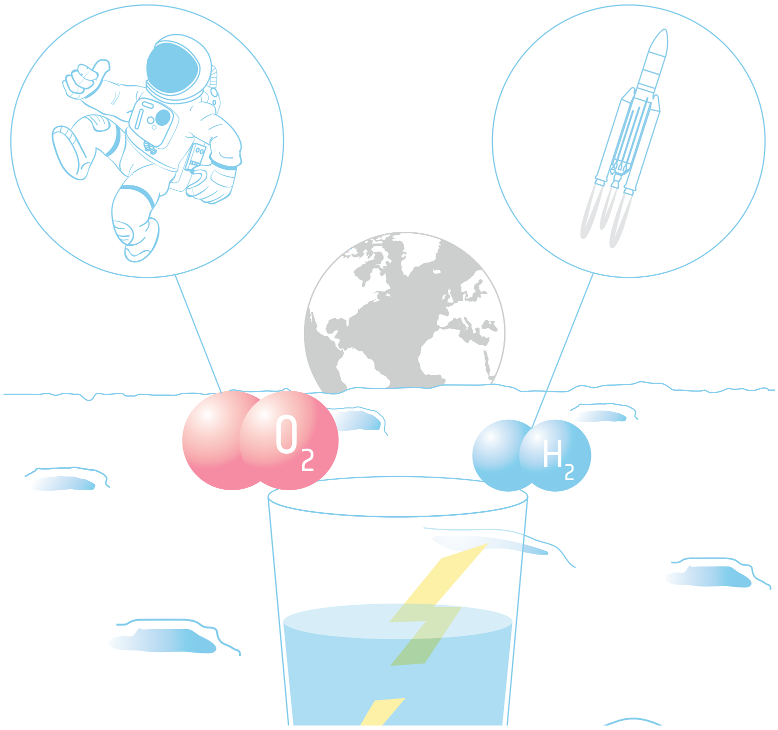
→ POWER FROM WATER

How to produce oxygen and hydrogen on the Moon



teacher guide & student worksheets

European Space Agency



Teacher guide

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→ POWER FROM WATER

How to produce oxygen and hydrogen on the Moon

Fast facts

Subject: Chemistry and Physics

Age range: 14-16 years old

Type: laboratory activity

Complexity: medium

Teacher preparation time: 1 hour

Lesson time required: 2 hours

Cost: medium (5-25) euros for activity 1 and 2 and high (50-100 euros) for activity 3

Location: laboratory

Includes the use of: Zinc and copper plates

Keywords: Chemistry, Physics, Moon, Electrochemistry, Volta Pile (battery), Electrolysis, Fuel cells

Brief description

In this set of three activities, students will learn about electrochemistry. In the first activity they will build a Volta Pile – a simple battery. This invention marked the beginning of electrochemistry. Students will then study electrolysis. Electrolysis uses electric current to split water into its components: Hydrogen and Oxygen. These products can be used as propellants for spacecraft and/or to provide Oxygen to support a crew. In the last activity, students examine and use a fuel cell.

Learning objectives

- Understanding how a battery works.
- Performing an experimental activity to confirm that certain chemical reactions can create electricity.
- Performing an experimental activity to confirm that electricity can make certain chemical reactions take place.
- Studying water electrolysis and its applications.
- Investigating fuel cells and their applications.
- Writing balanced equations for REDOX chemical reactions.
- Using equipment appropriately to make and record observations.

→ Summary of activities

Summary of activities					
	Title	Description	Outcome	Requirements	Time
1	Build your own battery	Building a Volta pile.	Introduction to electrochemistry; learning how a battery works.	None	45 minutes
2	Water electrolysis	Building an electrolyser and performing water electrolysis.	Learning about water electrolysis and its applications.	Completion of activity 1 is advised.	45 minutes
3	Fuel cell	Investigating a fuel cell.	Learning about fuel cells and their applications.	None. Completion of activity 2 is advised.	30 minutes

→ Introduction

Human exploration of the Moon requires resources: water, Oxygen, food, materials, propellant, etc. Bringing everything from Earth would be very inefficient and expensive, so instead mission designers are investigating how to use resources already available on the Moon. One of the most important resources is water. Scientists found evidence that water may exist in some areas towards the Moon's poles. On a future Moon mission, this water could be used to produce Hydrogen and Oxygen for propulsion and Oxygen for breathable air for the crew.

In this set of activities, we are going to investigate how to store energy in batteries and how to generate Oxygen and Hydrogen from water. In order to do so we need to learn about electrochemistry!

Electrochemistry is the branch of science that studies the relationship between electricity and chemistry. Certain chemical reactions can create electricity, as is the case in a battery. The opposite is also possible: electricity can make certain chemical reactions happen that wouldn't happen spontaneously.

In this resource students will be guided through the principles and chronology of electrochemistry, from the invention of the first battery (the Volta pile) to the modern use of fuel cells.

In this resource the students will build the following devices:

1. **Battery:** A device that generates electricity from chemical reactions.

2. **Electrolyser:** A device that uses electricity to make certain chemical reactions take place. In this case we will work with water electrolysis and break the bonds that hold the components of water molecules together

3. **Fuel Cell:** A device that produces electricity and heat from a chemical reaction.

→ Activity 1: Build your own battery

The Volta pile was the first battery ever invented by Alessandro Volta in 1799. Batteries generate electricity from chemical reactions, and the invention of the Volta pile marked the beginning of electrochemistry.

Batteries are often used on spacecraft as a means of electricity (power) storage and distribution. Traditional batteries contain all their usable energy and can only be discharged. Batteries used on space missions are often re-chargeable. They can be recharged with power from other sources, for example by solar energy. Batteries are crucial because they can provide electricity during periods with no access to other power sources (for example, when it is not in direct sunlight).

In this activity, students will construct a Volta pile – a simple battery from metal plates, dishcloth and vinegar. A Volta pile uses a spontaneous chemical reaction to create electricity.

Health and safety

The Volta piles should not be left connected in a closed container or unventilated room.

Equipment

- 6 zinc plates (per group)
- 6 cooper plates (per group)
- 1 dishcloth (per group)
- Scissors
- Vinegar
- Sandpaper
- 2 elastic rubber bands
- Wires with crocodile clips
- Multimeter
- (optional) AA Batteries

Figure 1

↑ Equipment needed for building a Volta Pile

Exercise

Start by introducing the concept of electrochemistry and the definition of electrical potential difference.

Building one Volta cell
Divide the class into groups of 3 to 4 students. The students should follow the instructions 1 and 2 in their student worksheet to assemble a simple Volta cell. After assembling the cell they should connect a multimeter and measure the electrical potential difference.

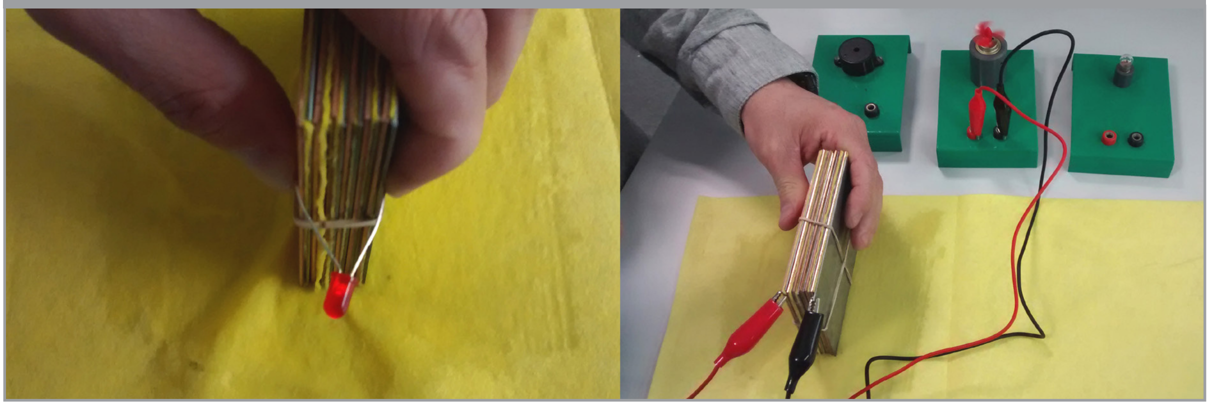
Ask the students to explain why they get an electrical potential difference and to explain what the function of each layer is in the Volta cell. Have the students write down the ionic equations of the reactions that occur in the Volta cell. Ask them to answer question 4 to 7 in their student worksheet.

Building a Volta pile
Now students should pile several Volta cells to get a Volta pile. Students should measure the electrical potential difference of the pile once per minute for 10 minutes and record the measurements in Table 1 in their student worksheets. In between measurements they should answer the related questions in their student worksheets.

Ask the students to plot the electrical potential difference of the Volta pile as a function of time. They should discover that the electrical potential difference decreases with time. Ask the students why they think this is happening.

To demonstrate that the Volta pile can generate current you can use it to light a LED or make a motor run as seen in Figure 2 and investigate how long the Volta pile is able to power the actuator.

Figure 2



↑ Making an LED light up using a Volta Pile (left) and powering a motor with a Volta Pile (right).

Ask the students to compare their Volta pile with a normal AA battery. Talk about how a normal battery works and what the limitations of the Volta Pile are. If time permits, let the students connect the battery and the Volta pile to different electrical appliances and measure the current in the circuit.

Results

These are the answers to the student worksheet activity 1:

3. From one Volta cell you should obtain approx. 1 V

4. The net ionic equation is:

$$Zn + 2h^{+} \rightarrow Zn^{2+} + H_2$$

The redox half reactions for the two half cells are:

OXIDATION (anode): $Zn \rightarrow Zn^{2+} + 2e^{-}$

REDUCTION (cathode): $2H^{+} + 2e^{-} \rightarrow H_2$

$Zn\,|Zn^{2+} \;||\; 2H^{+}|H_2|Cu$

(OXIDATION) || (REDUCTION)

(Electro-chemical chain notation)

The copper metal plate only serves as a "chemically inert" noble metallic conductor for the transport of electrons in the circuit and does not chemically participate in the reaction. The copper plate could be replaced by any sufficiently metallic conductor.

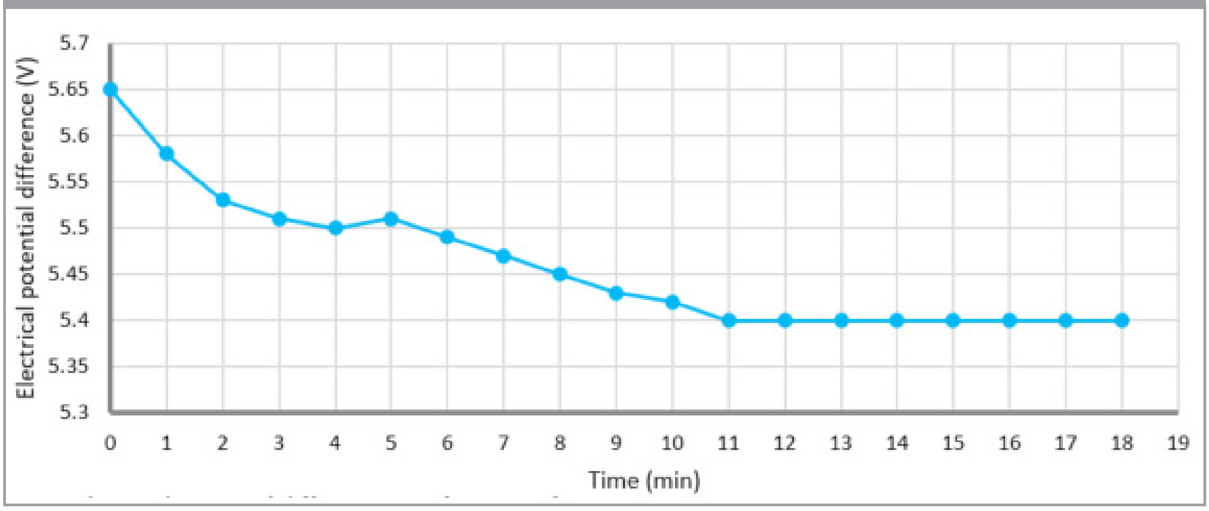
5. Zinc is oxidised and releases electrons (anode). Hydrogen (hydronium = H_3O^{+}) is reduced and gains electrons (cathode).

6. An oxidised layer accumulates on the plate's surface (which makes the metal look dull). This decreases the useful area for ion exchange. Sanding removes the oxidised layer.

7. Vinegar is an electrolyte. It allows the exchange of ions between the plates and increases the electrical potential difference. All acids, like vinegar, release H^{+} ions that are necessary for the reaction. Even water releases H^{+} ions, but in a very low quantity. Adding a salt or an acid increases conductivity. We could use any other substance that acts as an electrolyte (salt or acid)

10. Example result obtained from a Volta pile consisting of six Volta cells

Figure 3



↑ Electrical potential difference as a function of time.

11. The voltage drops with time because the internal resistance of pile increases. The zinc surface is oxidized, which results in the reaction surface area being reduced. Vinegar (and other acids) also cause oxidation. Furthermore, in the Volta pile, hydrogen bubbles are accumulating on the Copper surface (polarization). Commercial batteries use very different materials, which oxidise far less than zinc. Some types of batteries use a substance that removes the accumulated hydrogen or vents it off. For these reasons, commercial batteries can continue operating much longer.

12. An AA battery usually has an electrical potential difference of 1.5 V unless otherwise specified by the manufacturer. In our example we obtain 1 V from one Volta cell and we obtain 5.5 V with the Volta pile (six Volta cells). When we increase the number of cells, we increase the total contact surface for ion exchange. The limiting factor to the current output is the internal resistance (which is high).

13. Batteries would be useful for Moon exploration as a way to store energy. Batteries used on space missions are often re-chargeable from other sources, for example by solar energy. Batteries are crucial because they can provide electricity during periods without access to any other power source (for example, when there is no direct sunlight). Bringing non-rechargeable batteries as the only power supply would be both heavy, inefficient and not sustainable.

Discussion

Discuss with the students the importance of the invention of the Volta pile. How would our lives be without batteries? Could we design a limitless battery? Discuss the reasons for inefficiencies: weight and limited lifetime versus storing capability and power output.

Talk about how energy can only be transformed, not lost or created. Discuss why we still talk about losing energy (because heat is a difficult form of energy to use).

The Volta pile can be reset by cleaning the metal plates with sandpaper and soak the dishcloth with the electrolyte again. Discuss if any rechargeable battery can be recharged indefinitely.

→ Activity 2: Electrolysis

Electrolysis uses electricity to make chemical reactions occur that would not happen spontaneously. In this activity, students will build an electrolyser: a device that introduces an electric current into a liquid using two electrodes. They will use the device for water electrolysis and discover that it is possible to split water into its components: Oxygen and Hydrogen.

Health and safety

When testing the gases be sure to keep a safe distance and use long matches or splits to avoid being burned.

Equipment (per group)

- Plastic container with lid (with two holes – see preparation)
- 2 test tubes
- 2 steel push pins
- 2 beakers
- Copper wire
- Battery (optional: a solar cell)
- 400 cm^3 distilled water + 12 g NaOH (3% dis-solution)
- Distilled water
- Gloves

Figure 4

↑ Equipment needed for building an electrolyser

Preparation

Drill two small holes in the bottom of the container and two holes in the lid (with the diameter of the test tubes) as shown in Figure 5.

Figure 5

↑ Preparation of box for electrolyser

Exercise 1

Have the students balance the global equation for the water electrolysis process. Then help them write and understand the oxidation and reduction reactions.

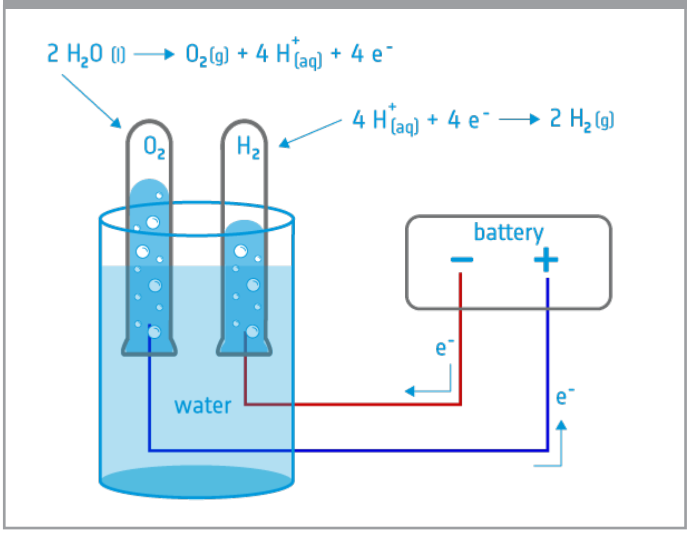
An illustration of the electrolysis setup can be seen in Figure 6.

Let them build their own electrolyser following the instructions in Annex 1. Remind them to time the electrolysis process in order to later calculate the production rate of Oxygen. Ask the students to answer question 5 to 7 in their student worksheet about the electrolysis process.

Relate the experiment with Oxygen production for space missions. Ask the students to answer questions 8 and 9 in their student worksheet to investigate if their electrolyser could produce enough Oxygen for astronauts on the Moon.

If you wish to extend the exercise, students can do the experiment with distilled water, tap water (containing salts) and water with electrolyte.

Figure 6



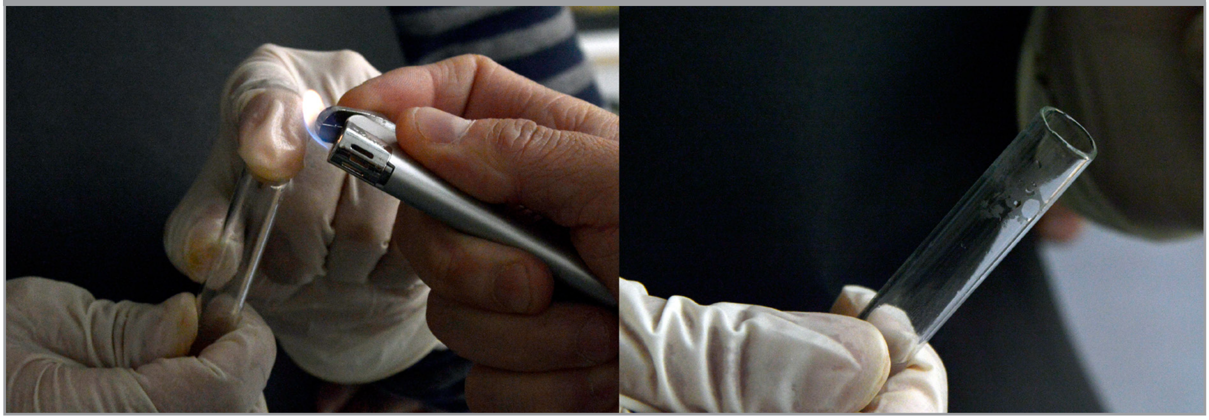
↑ Illustration of the electrolysis setup and process

Identifying the gases

You can do this test as a classroom demonstration or let the students test it themselves. Before performing the test, ask the students if they have any ideas how they could test the gases in the tubes.

Hydrogen test: Place a finger on the opening of the test tube to prevent the Hydrogen from escaping, then turn the test tube around and place a long match (or lighter) at the opening of the tube (see Figure 7). You should hear a squeaky pop which confirms the presence of Hydrogen (the sound indicates a mini explosion). The reaction $2\ H_2(g) + O_2(g) \rightarrow 2H_2O(L)$ will produce a small amount of water at the bottom of the test tube.

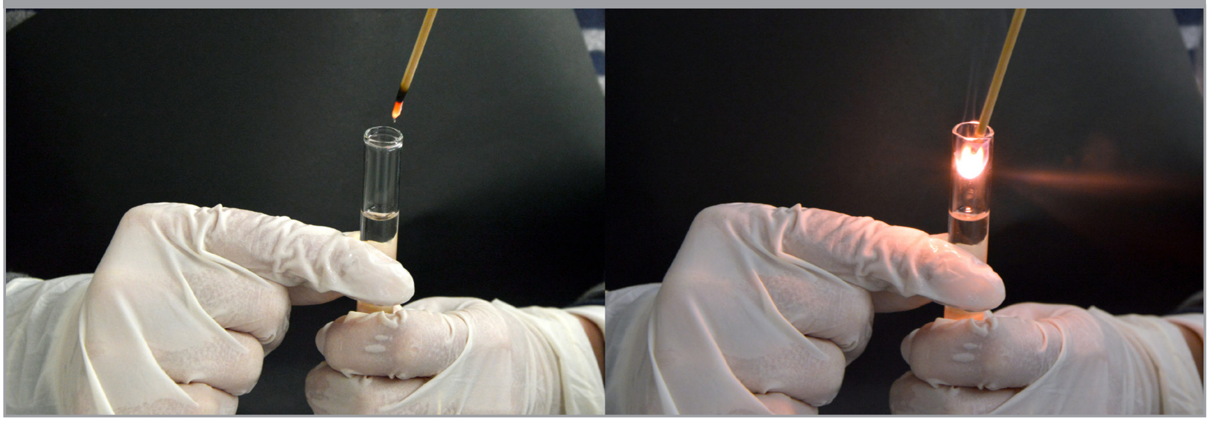
Figure 7



↑ Testing for Hydrogen with lighter (left) and presence of water in the test tube (right)

Oxygen test: Place a finger on the opening of the test tube to prevent the Oxygen from escaping, then turn the test tube around and place a lit splint at the opening of the tube (see Figure 8). The presence of Oxygen will reignite the stick.

Figure 8



↑ Testing for Oxygen with lit splint (left). Splint re-ignites in the presence of pure oxygen (right)

Results

These are the answers to the student worksheet activity 2:

1. The global reaction of water electrolysis:
   $2H_2O\ (l) \rightarrow 2H_2\ (g) + O_2\ (g)$

2. Anode oxidation:
   $2H_2O\ (l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

3. Cathode reduction:
   $4H^+(aq) + 4e^- \rightarrow 2\ H_2(g)$

6. NaOH is an electrolyte. Adding an electrolyte speeds up electrolysis because it increases the electrical conductivity of water (i.e. deceases resistance to electricity). Salt, acid or base could be used as an electrolyte. For the specific case of alkaline water electrolysis, a strong base like sodium hydroxide (or potassium hydroxide) is used as the electrolyte, thus avoiding the corrosion problems caused by acid electrolytes (corrosion of metallic electrodes) .

7. In the equation $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$ the products are two Hydrogen atoms for each Oxygen atom. For that reason, we observe twice the amount of Hydrogen than Oxygen.

8. Using a Volta pile of six Volta cells with an electrical potential difference of 6 V, it is possible to produce 3ml O_2 in 4 hours = 18 ml/day of molecular Oxygen (O_2).

9. Obtaining 18 ml/day is equivalent to $1.8*10^{-5}\ m^3$/day. We can use the ideal gas law to calculate the number of mole of O_2, and from that the mass:

$$P * V = n * R * T$$

$$n = \frac{P*V}{R*T} = \frac{101325pa*1.8*10^{-5}m^3}{8.314\frac{m^3*pa}{K*mol}*293\ K} = 7.48*10^{-4}mol$$

$$m = n*M = 7.48*10^{-4}mol*32\frac{g}{mol} = 0.0239g$$

   This is the same as $2.4 * 10^{-5}$ kg and hence only supplies:

$$\frac{2.4*10^{-5}kg}{0.84kg}*100\% = 0.0028\%$$

   of the amount needed by one astronaut per day.

10. We could make the production go faster by increasing the electrolyte concentration (in this case the concentration of NaOH) or by using a more powerful battery.

11. Oxygen is vital in order for the astronauts to breathe on the Moon. Hydrogen in combination with an oxidant (for example Oxygen) can be used as a source of fuel to propel spacecraft going further into space or to propel rovers to explore the lunar surface and expand human presence. If we had to bring the Oxygen and Hydrogen we would need containers and a lot of mass and volume. This would be very expensive. In order to make an outpost on the Moon sustainable we need to recycle as much as possible and transform waste products (for example CO_2, urine, sweat, food waste, metabolic waste etc.) into O_2 and water that can be used again. This is what Life Support Systems are designed to do and technologies for optimizing recycling are being tested on the International Space Station. However, to be more independent from Earth re-suppling the outpost we also have to learn to produce most of what we need on the Moon (in-situ resource utilization).

→ Activity 3: Fuel cell

In this activity, students will use the products of water electrolysis (H_2 and O_2) in a fuel cell. They will investigate how fuel cells produces electricity and heat from a chemical reaction. Students will consider the possibilities and limitations of fuel cells for Moon exploration.

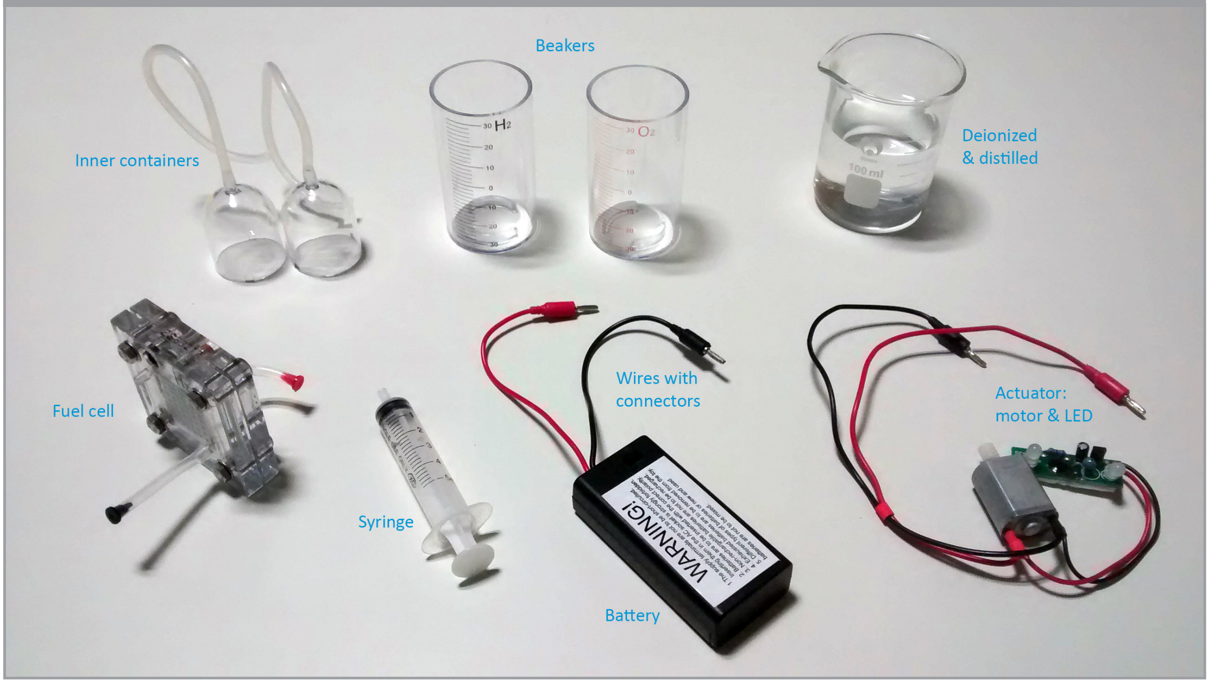
This exercise can be done either as a demonstration or if time permits as a hands-on activity. The experiment requires a fuel cell, which can be acquired online[1].

Equipment

A Fuel cell car science kit[1] or:

- A fuel cell
- A syringe
- Deionized & distilled water
- Power supply (battery, solar cell)
- Silicon tubing and caps
- 30 ml beakers and inner containers (see annex 2)
- Wires with connectors
- Actuators (Motor, LEDs, car etc.)

Figure 9



↑ Equipment needed for building a fuel cell system.

Exercise

Start by introducing fuel cells to the students. Fuels cells builds on the knowledge of water electrolysis and the Volta pile and offers an opportunity to wrap up the resource.

Ask the students to follow the instructions in annex 2 (or prepare the experiment beforehand for a demonstration). Ask the students to fill in questions 1 to 5 in their student worksheets and reflect on the advantages and limitations of fuel cells.

[1] This resource uses the Fuel cell car science kit that can be bought online http://www.horizonfuelcellshop.com/europe/product/fuel-cell-car-science-kit/.

Results

1. Overall reaction in the fuel cell:

$$2H^2(g) + O_2(g) \rightarrow 2H_2O(l) + \text{electric energy} + \text{(waste) heat}$$

2. Anode and cathode reactions

Anode:	Cathode:
$H_2(g) \rightarrow 2H^+(aq) + 2e^-$	$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O\ (l)$

3. If Oxygen and Hydrogen is readily available and requires little energy to obtain, then a fuel cell is a cheap and clean energy source. It does not pollute: it only produces water and energy. This makes it a potential solution for clean energy on Earth. If we need to produce H_2 and O_2 with electrolysis before we can run the fuel cell, or if H_2 and O_2 are valuable and limited, then fuel cells might not be the optimal solution. We will lose energy since we are need another power source in the first place to create the components for the reaction. However, if we combine fuel cell technology with a renewable energy source (for example solar energy) then we can use fuel cells without polluting.

4. Both the Volta pile, the electrolysis and the fuel cell are examples of electrochemistry in practise. A Volta pile uses a chemical reaction to generate electricity, water electrolysis uses electricity to make a chemical reaction happen that would not otherwise occur, and a fuel cell reverses the electrolysis reaction and generates electricity again from the products of the electrolysis.

→ POWER FROM WATER

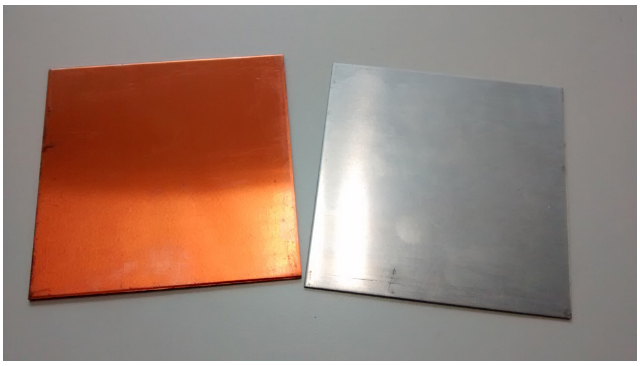
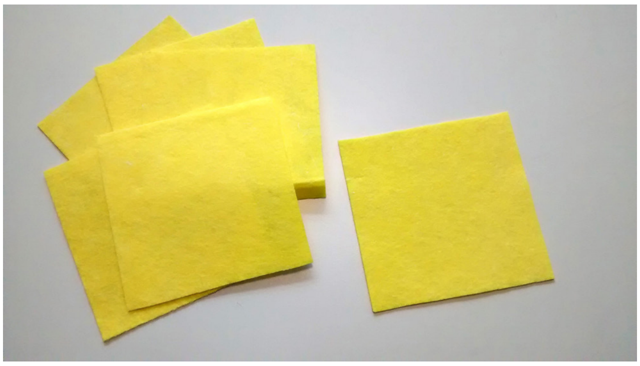
How to produce oxygen and hydrogen on the Moon

→ Activity 1: Build your own battery

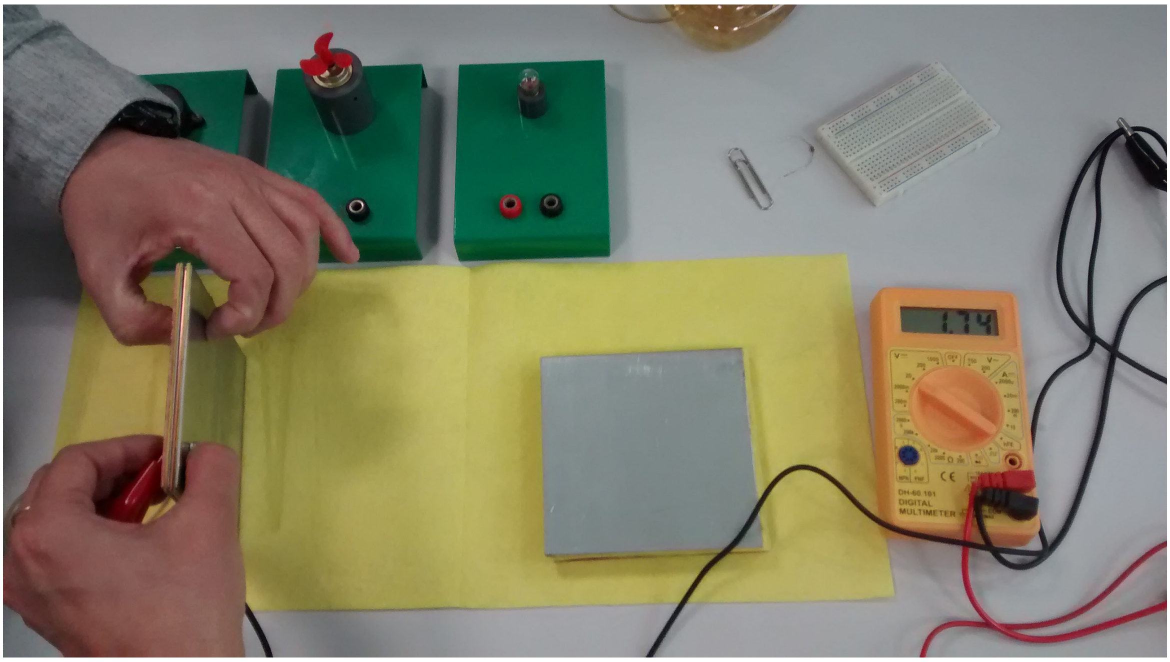
In 1799 Alessandro Volta, one of the pioneers of electricity and power, invented the first battery: the “Volta pile”. With this invention, he proved that certain chemical reactions can create electricity. Now you will follow in his footsteps and build your own Volta pile.

Exercise 1

1. Sand both sides of the zinc and copper plates with sandpaper and cut 6 squares from the dishcloth in the same size as the plates.

2. Place one dishcloth square on top of a zinc plate and soak it in vinegar. Then place one copper plate on top of the dishcloth. You now have one Volta cell. Connect the wires with the crocodile clips to the first and last plate, then connect to a multimeter.



3. What is the electrical potential difference from one Volta cell? ____________ V

4. Write down the net ionic equation, which shows the overall process occurring in the cell:

5. Which component is oxidised and which one is reduced in this reaction?

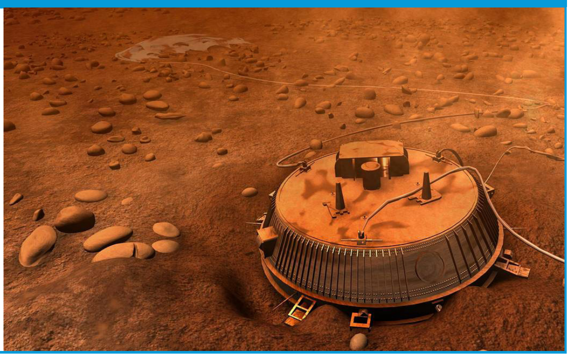
6. Why is it a good idea to sand the metal plates?

7. Why do we soak the cloth in vinegar? Could we use another substance? Explain.

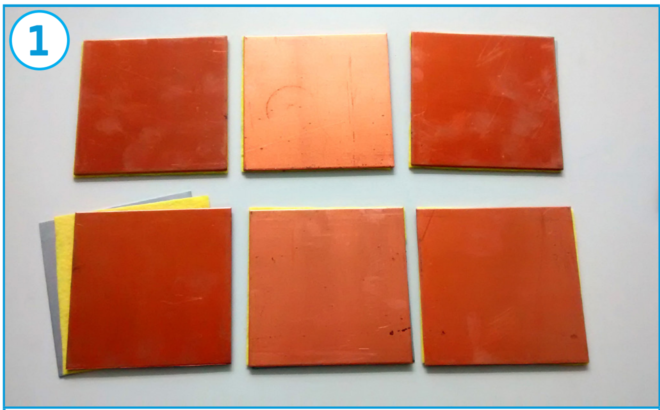
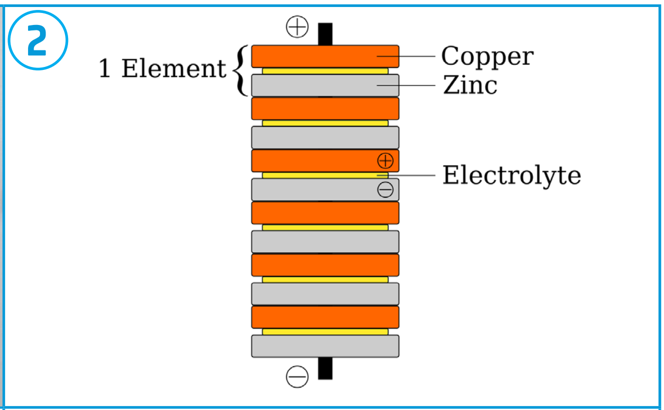
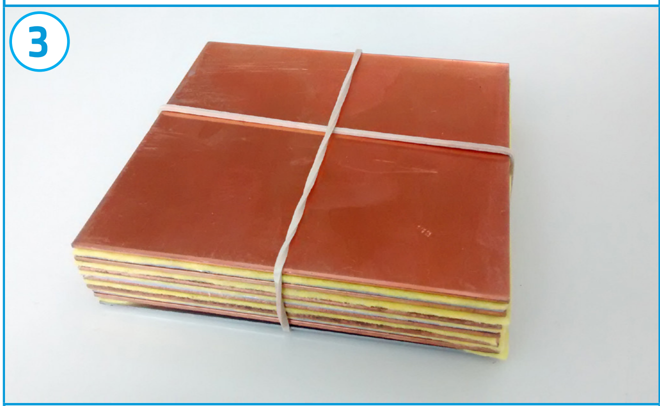
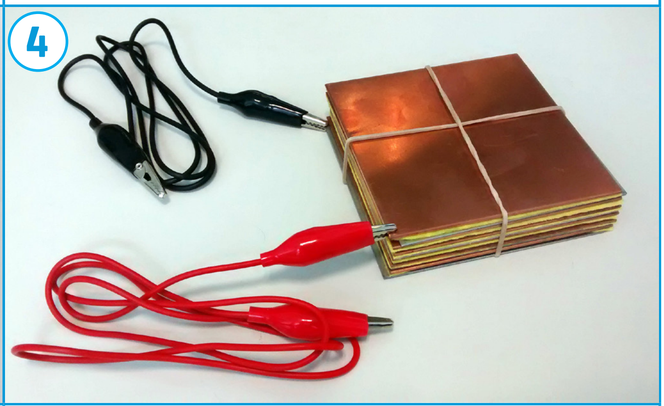
Did you know?

ESA's Huygens probe, which landed on the surface of Titan (Saturn's largest moon), relied on non-rechargeable lithium sulphur dioxide batteries. These were chosen because they could be left inactive during the seven-year trip to Saturn, but still retain sufficient capacity for the landing on Titan.

Artist impression of the Huygens probe on the surface of Titan.→



8. Follow the instruction below to build a Volta pile.

1 	2 
Create six Volta cells following the instructions from exercise 1	Stack the cells in the order shown
3 	4 
Tie elastic rubber bands around the pile to keep the pile in place.	Connect the wires with the crocodile clips to the first and last plate and connect to a multimeter.

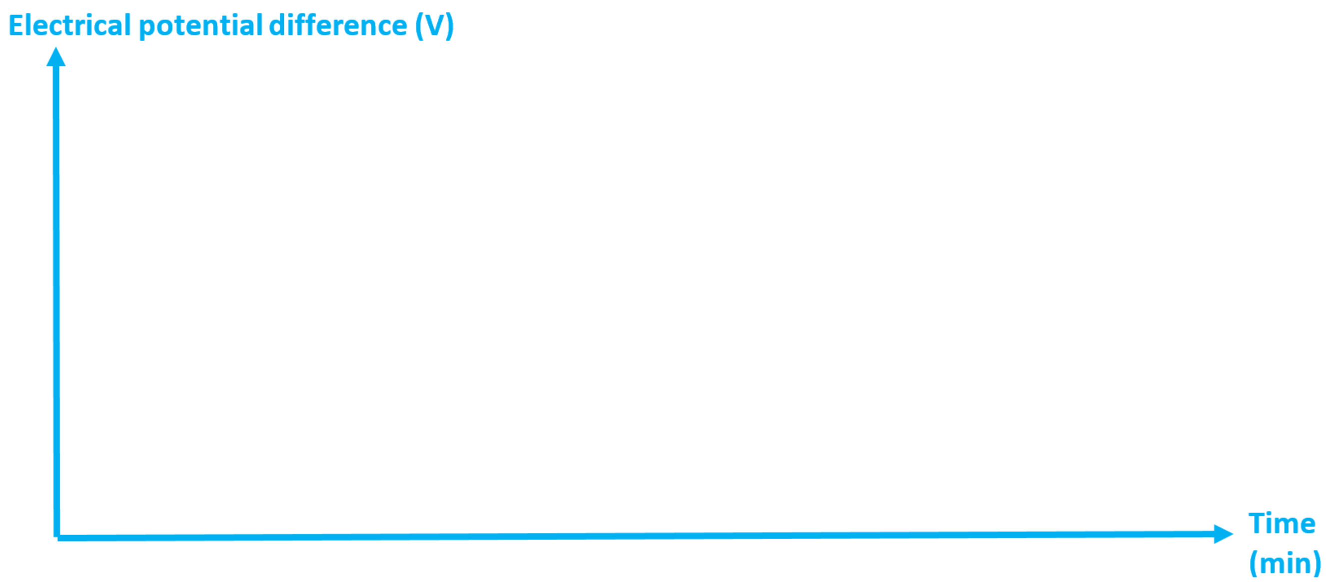
9. Measure the electrical potential difference of the Volta cell immediately after you have finished assembling it. Take measurements every minute for 10 minutes and record your measurements in Table 1.

Table 1

Time (min)	Electrical potential difference (V)
1	
2	
3	
4	
5	
6	
7	
8	
9	
10	

↑ Recordings of electrical potential difference for 10 minutes.

10. Draw a graph of electrical potential difference as a function of time.



11. What do you observe? Explain.

12. What is the electrical potential difference of the pile compared to a normal AA battery?

13. How could batteries be used for Moon exploration? What are the advantages and disadvantages of batteries?

→ Activity 2: Water electrolysis

Exercise

1. Balance the global reaction below:

$$\square H_2O(l) \rightarrow \square H_2(g) + \square O_2(g)$$

2. Complete the reaction equation of the anode oxidation:

$H_2O(l) \rightarrow$

3. Complete the reaction equation of the cathode reduction:

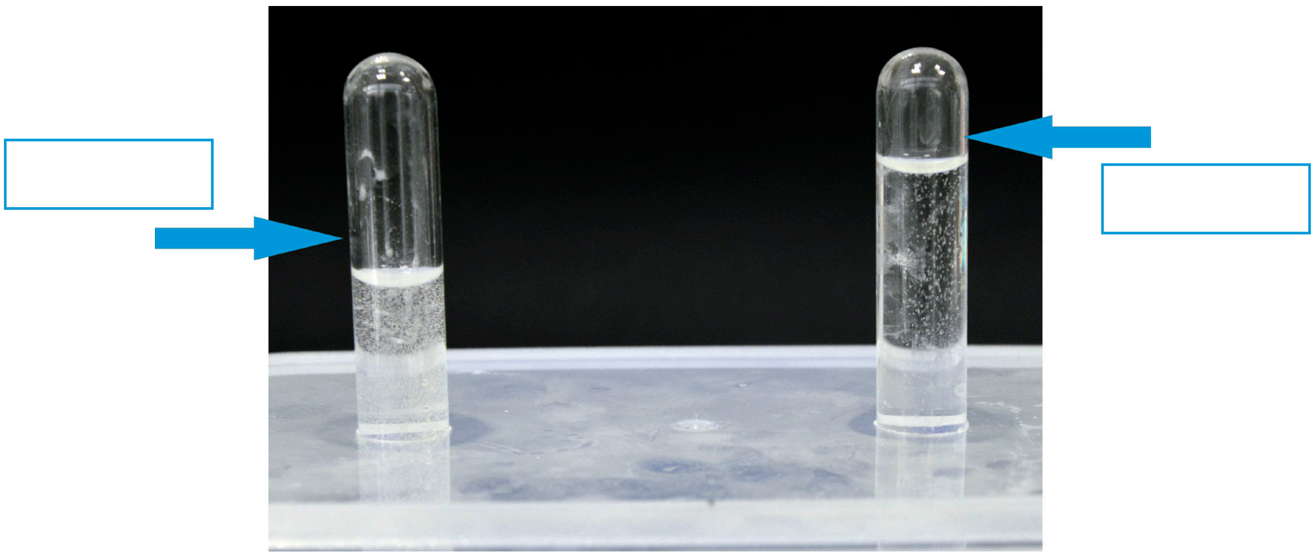
$\rightarrow H_2(g)$

4. Build your own electrolyser by following the steps in Annex 1. When starting the electrolysis make sure to start a stopwatch too. You will need the time of the electrolysis to calculate the Oxygen production.

5. Describe what happens during the electrolysis.

6. What is the function of dissolving NaOH in the water?

7. Which of the tubes contains oxygen and which hydrogen? How do you know?



8. Measure how much Oxygen you have produced and calculate your production rate per minute.

9. Assume that one astronaut breathes 0.84 kg of molecular Oxygen (O_2) per day. Can your electrolysis system supply the required Oxygen?

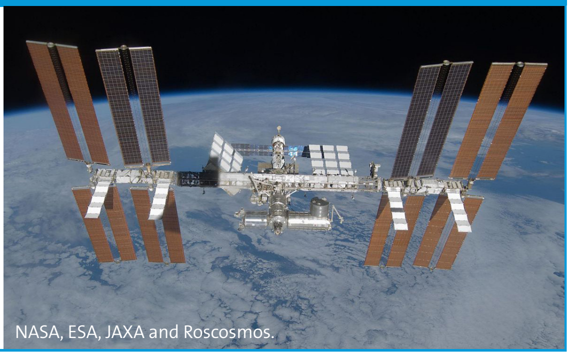
10. How can you produce more Oxygen per day?

11. What are the advantages and disadvantages of using water electrolysis on the Moon?

Did you know?

Electrolysis of water is the main method of generating Oxygen on board the International Space Station (ISS). Water is collected from urine, wastewater, and condensation and split into Oxygen and Hydrogen in the Oxygen Generation System (OGS). The station's football-field-sized solar arrays are the power source. A similar system could be used on the Moon.

The international Space Station (ISS) orbiting Earth →


NASA, ESA, JAXA and Roscosmos.

→ Activity 3: Fuel cell

Fuel cells can be used to power rovers or space ships or anything in between. Fuel cells do the opposite of electrolysis: They combine H_2 and O_2 and produce H_2o(water) and energy.

After observing how a fuel cell works, answer the questions below:

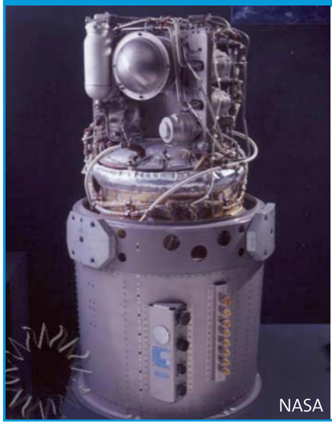
1. Balance the global reaction:

$$\square H_2(g) + \square O_2(g) \rightarrow \square H_2O(l) + \text{energy} + \text{heat}$$

2. Write the chemical reactions occurring at the anode and the cathode:

3. What are the advantages and disadvantages of fuel cells? How can we use them for missions to the Moon?

Did you know?



Fuel cells were the primary source of electric power on the Apollo programme that landed humans on the Moon. Apollo fuel cells used Oxygen and Hydrogen, stored as liquids at extremely cold temperatures, that when combined chemically yielded electric power and water for drinking. Fuel cells were also used on the Space Shuttle fleet as one component of the electrical power system. The space shuttle fleet flew 135 missions between 1981 and 2011.

← Apollo fuel cell.

4. What are the differences and similarities between the Volta pile, electrolysis and the fuel cell?

→ Links

ESA resources

Moon Camp Challenge
esa.int/Education/Moon_Camp

Moon animations about the basics of living on the Moon.
esa.int/Education/Moon_Camp/The_basics_of_living

ESA classroom resources:
esa.int/Education/Classroom_resources

Extra information

Voltaic cell:
youtube.com/watch?v=9OVtk6G2TnQ

Electrolysis:
youtube.com/watch?v=dRtSjJCKkIo

Fuel cells:
youtube.com/watch?v=OmVnIIgDA7o

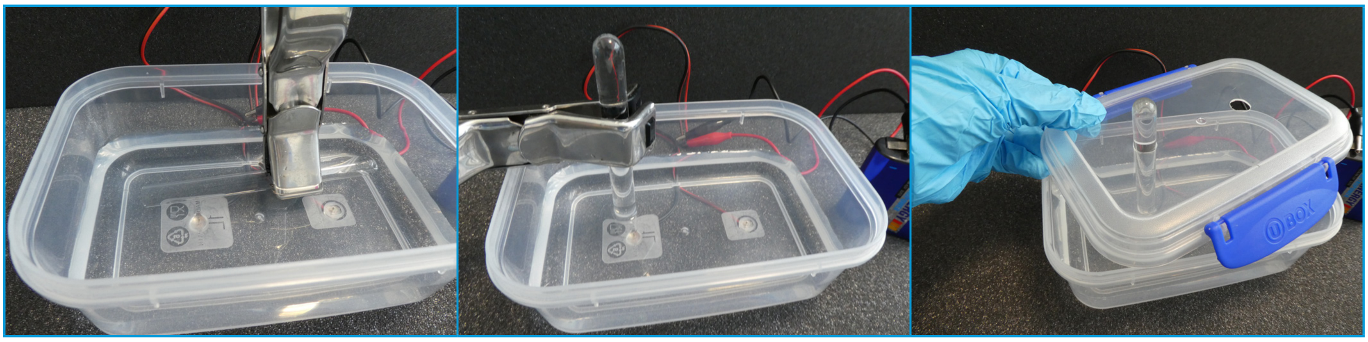
Fuel cell car science kit:
www.horizonfuelcellshop.com/europe/product/fuel-cell-car-science-kit

→ Annex 1: Electrolyser

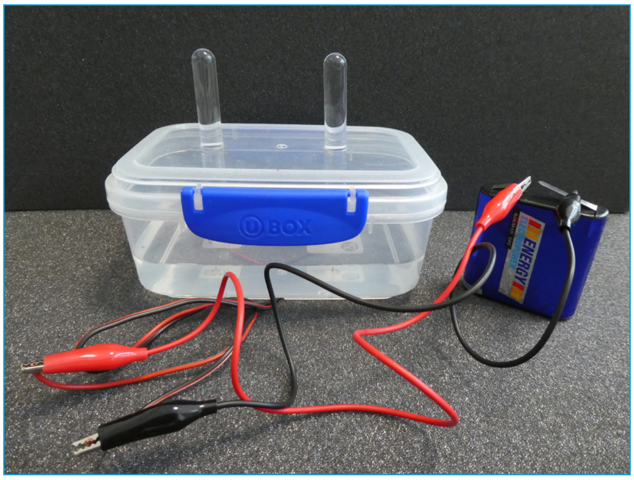
1. Use push pins to fasten the wires to the box

2. Dissolve NaOH in water (3%) and fill the container

3. Raise the test tubes over the punch pins (make sure they stay filled with water). Place the lid on top to keep the tubes in place

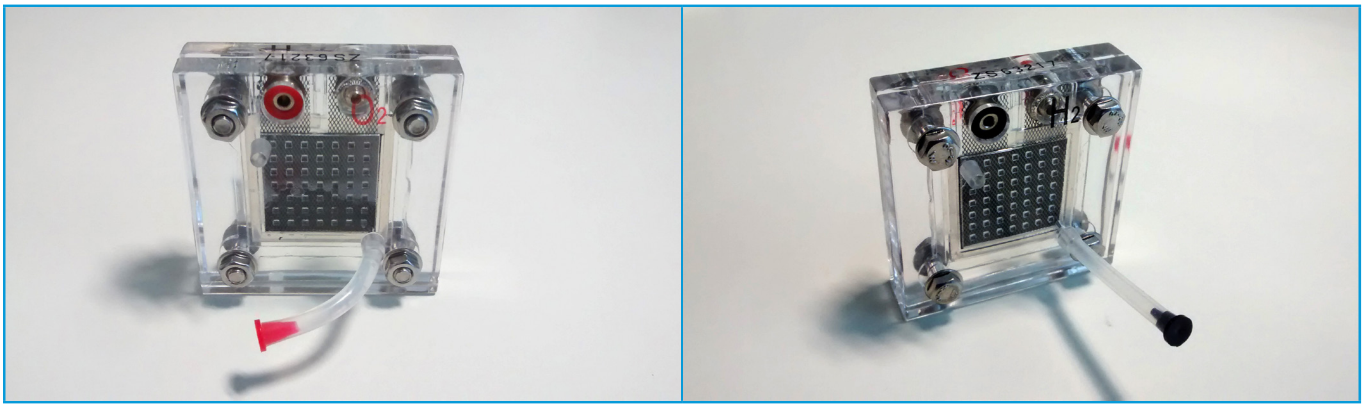


4. Connect the power source and start the electrolysis. Start a timer and record how long the process takes.



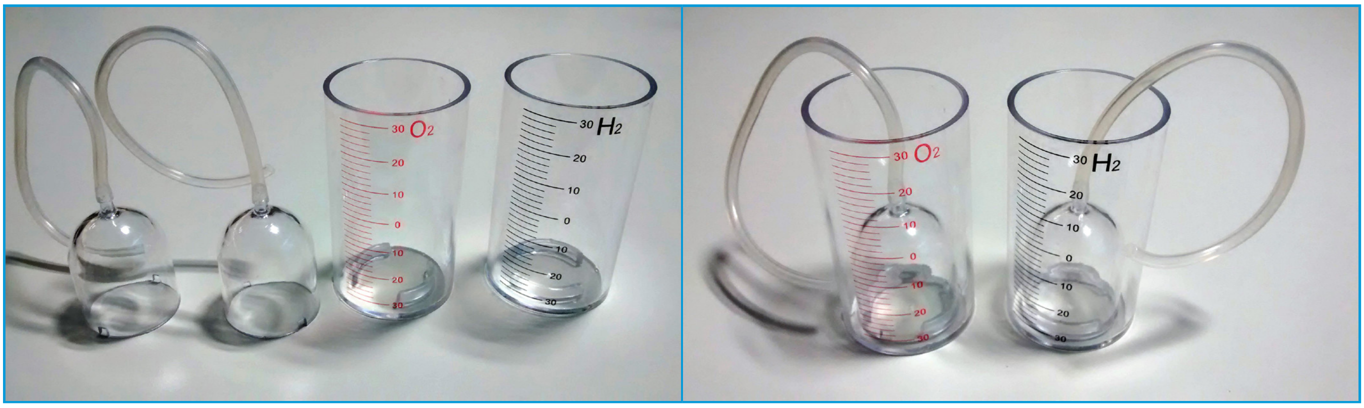
→ Annex 2 – Fuel cell

1. Connect two silicon tubes (4cm) to the nozzles on the fuel cell located on the lower section on both sides of the fuel cell.

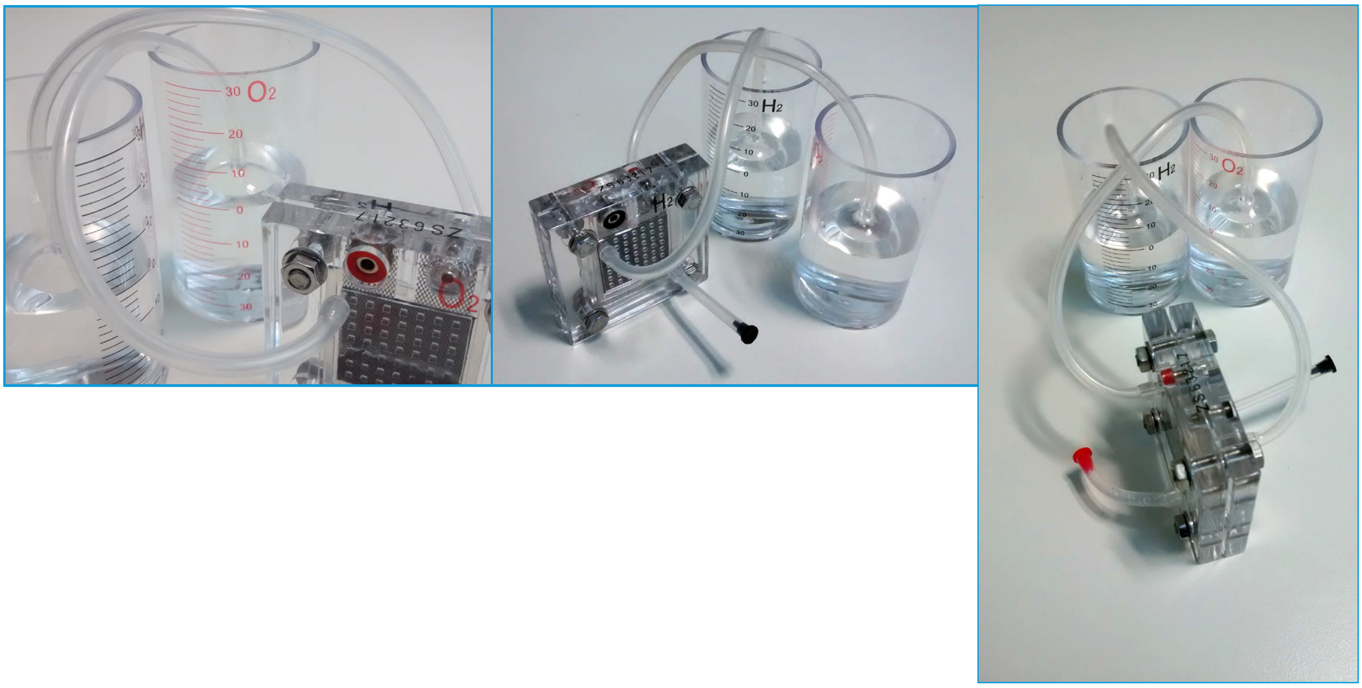


2. Remove the red cap from the tube connected to the oxygen side of the fuel cell. Using the syringe, inject distilled water into the fuel cell until the fuel cell is half filled with water

3. Place the dome shaped inner containers into the beakers and add water up to the "0" ml (the water should almost cover the inner container). Make sure there is no air trapped inside the inner containers.

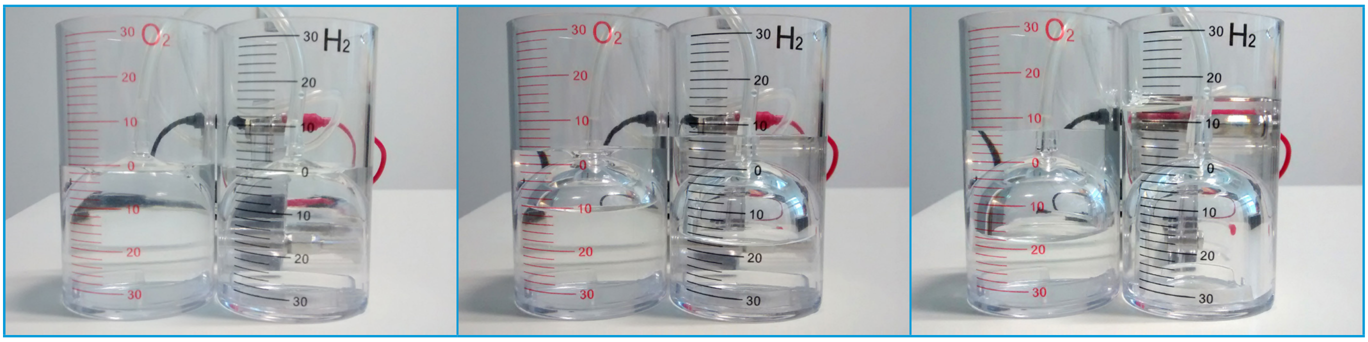


4. Connect the silicon tubes to the upper nozzles of both the O_2 and H_2 side of the fuel cell.



5. Connect the battery pack as seen below. Switch the battery pack to the "on" position and electrolysis should begin.

6. Observe how H_2 and O_2 start to fill the inner containers. When H_2 begins to bubble out of the container the tank is full.



5. Disconnect the battery and connect an actuator (car, motor, LED etc.). H_2 and O_2 combines again producing H_2o and energy to power the motor.