

Return to 1616 education

Commemorating the 400th anniversary (2016) of Dirk Hartog's landing on the Western Australian coast

A Year 9 Science Education Resource - Year 9 Chemical Science

Ships, Shipwrecks and Reefs

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De Eendracht

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GOVERNMENT OF
WESTERN AUSTRALIA



Ships, Shipwrecks and Reefs

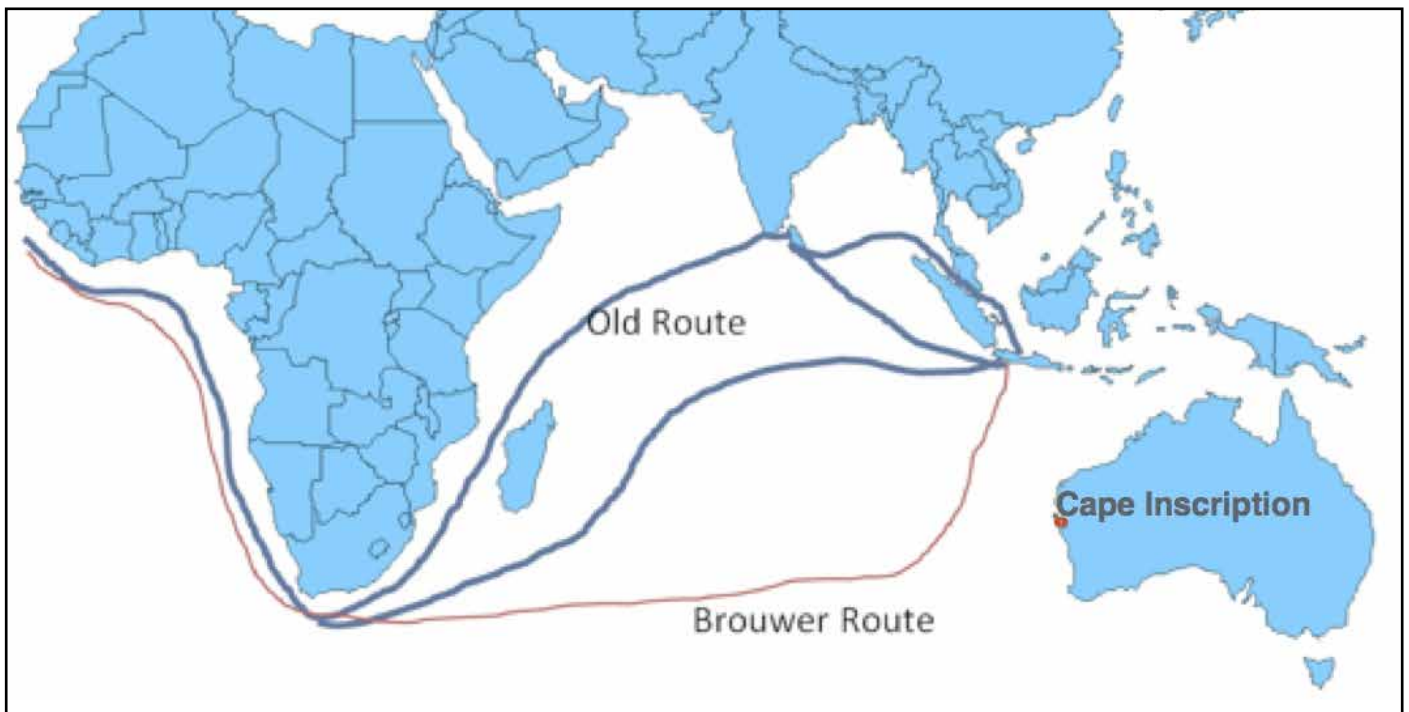
Dirk Hartog Island

In 1616, Dutch skipper Dirk Hartog in the ship Eendracht, accidentally encountered the west coast of Shark Bay at what is now known as Dirk Hartog Island. At Cape Inscription, Hartog's crew left a flattened pewter plate nailed to a wooden post and inscribed with a record of his visit, which now represents the oldest physical evidence of European contact with Australia. He was one of many sea captains in what has become known as the Age of Exploration.

Hartog's voyage began on the 23rd of January 1616 to transport goods between the Dutch port of Texel and Batavia (Jakarta) in the Dutch East Indies.

The Eendracht sailed in a convoy of other ships but became separated from them in bad weather and Hartog decided to travel to Batavia alone.

Some years earlier Dutch navigator Hendrik Brouwer had discovered that by sailing a southerly course, at latitude 40oS ships would make a high-speed passage due to the strong westerly winds known as the "Roaring Forties". The Dutch East India Company (VOC) ordered its captains to take this new route and hence reduce the overall voyage by six months.



The old and new routes

Dirk Hartog unexpectedly came upon Cape Inscription on 25 October 1616 at a latitude around 26° South. He had sailed too far east and had not turned north early enough as shown above on the Brouwer route.

The new islands he found were unknown to Europeans, and Hartog's ship was only the second European vessel to visit Australia. A decade earlier, Willem Janszoon, who captained the Duyfken, had landed on Cape York Peninsula (now in Queensland).

Hartog and crew landed on Dirk Hartog Island and found it to be uninhabited. Hartog spent three days there, finding nothing of great interest or value to him or his company.

Activity 1 - Teacher Demonstration: Analysis of tamala limestone

Dirk Hartog Island: Geology

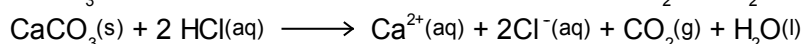
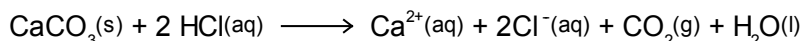
During the Pleistocene Ice Age (125 000 – 10 000 years ago), the sea level was over 100 metres lower than it is today.

This fall in sea level exposed large amounts of shell and coral. Strong winds from the south eroded the shell and coral into sand sized particles and blew them into dunes all along the West Australian coast from Albany to Shark Bay. The dunes also contained sand that had been formed by the ocean eroding the coastline during the previous million years. The mixture later formed into rock when the lime (calcium carbonate) content dissolved to cement the grains together.

Today we call this rock the Tamala Limestone and Dirk Hartog Island is the western-most deposit. It is a soft easily cut limestone and was used extensively in the convict buildings such as the Fremantle Prison, the Round House and other major buildings of early Western Australia.

The Tamala limestone is formed from wind-blown shell and coral fragments (calcium and magnesium carbonates) mixed with sand. In this experiment, you will analyse a sample of Tamala limestone to find out how much sand it contains.

The first task is to weigh a limestone sample, then react it with hydrochloric acid until the reaction ceases. The sand (called silica) will be left unreacted in the beaker.



The mixture is then filtered to remove the un-reacted solids.

Equipment required

- Limestone rock sample, about 2 g
- Hydrochloric acid [HCl] 6.0 mol L⁻¹ (15 mL)
- Two 250 mL beakers
- Drying oven
- Bunsen, tripod and gauze mat
- Filter paper
- Filter funnel

Procedure

IMPORTANT: Have students record all observations as you work and record all data as required.

1. Obtain a sample of limestone rock having a mass of about 2 g. Weigh the sample and record its mass. Place the limestone in a 250 mL beaker.
2. Obtain 10 mL of 6 M hydrochloric acid (CARE: very corrosive!). Add the hydrochloric acid to the limestone and, using low heat, heat for 5 minutes or until the bubble-producing reaction ceases. CARE: If the mixture is almost dry before the bubbling stops, allow it to cool, then carefully add another 3 mL of 6 M hydrochloric acid to allow the reaction to go to completion.
3. Weigh and then fold a filter paper and set up a filter funnel. Filter the sand/acid mixture to retrieve the unreacted sand. Wash out the beaker to ensure all solids have been removed.
4. Remove the filter paper and allow it to dry overnight.
5. Next day weigh the dry filter paper containing the sand.

Processing of results and questions

1. Calculate the mass of the dried sand.
2. Calculate the percentage purity of your limestone sample as follows:

$$\text{Percentage of sand} = \frac{\text{Weight of sand}}{\text{original weight of limestone}} \times 100$$

Wooden Ships

From the first time people set out to sea their vessels have been made from many materials. The earliest were dug-out canoes. As ships were made larger to carry more cargo and weather larger seas, the materials used in their construction had to become stronger and be longer lasting.

When wood is in contact with seawater in jetties or ships it is liable to be attacked by various marine organisms. Some, like seaweed, form massive growths that slow the ship's travel through the water while others, like teredo worms eat wood much like termites do on land



Figure 1: Part of the Busselton Jetty destroyed by Tereido worms.

Many methods have been tried to slow this attack. Lead and copper sheathing was used on wooden ships in the latter 18th century and afterwards, as a method of preventing this damage.

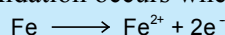
The British pioneered the use of copper sheathing and the first ship with this protection sailed in 1763.

Vessels made from iron and steel are stronger than wooden vessels and are not attacked by marine organisms but can corrode rapidly in the saline environment. This corrosion is the chemistry we shall explain in this unit.

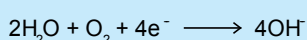
Electrolytes

The ocean contains dissolved salts that enable it to conduct electricity. These salts are called electrolytes. Most of the ocean salts originated on land and were washed into the ocean by rainwater while some of the salts originated in hydrothermal vents on mid-ocean ridges and volcanic hot-spots. Because the ocean will conduct electricity special types of chemical reaction that involve electrons being lost and gained by various substances can occur. These reactions are called oxidation-reduction reactions (redox reactions).

Oxidation occurs when a substance loses electrons. An example is:



While reduction is the process of gaining electrons:



Oxidation reactions and reduction reactions always occur together. Oxidation-reduction reactions can occur when ions are free to move in solid and liquid electrolytes.

Redox chemistry was first studied by Luigi Galvani in the 1780s. He accidentally discovered that by touching a dissected frog's leg using two different metals the leg would move. He believed that life and movement were caused by animal electricity. It was because of this discovery that Mary Shelly wrote her famous story, Frankenstein.

Alessandro Volta also studied electricity and did not believe in animal electricity. He developed the first electric battery made from zinc and copper plates separated with pieces of cardboard while the whole was immersed in sulfuric acid.

Volta was a true scientist;

“You must be ready to give up even the most attractive ideas when experiment shows them to be wrong.” Alessandro Volta

He developed the activity series of metals. He determined that if instead of copper and zinc he used metals further apart he would generate more electricity.

Activity Series of Metals
Zinc
Iron
Tin
Lead
Copper
Silver
Gold
Graphite

Activity 2: Research Investigation

Metal protection in marine environments:

Wooden planked boats were constructed by the:

- Phoenicians using copper alloy spikes, nails and fasteners
- Greeks and Romans using iron to replace the weaker copper
- Vikings who used iron nails to secure overlapping shingles to cover oak planks

Lead, and then copper, sheeting was fixed to the hull of ships to prevent biofouling by marine organisms. The sheeting was fixed initially using iron nails but these were found to rust. Copper nails were used in their place.

A bronze alloy (60% copper and 40% zinc) was developed to replace the copper sheeting. This alloy was found to be stronger and could be used to make thinner sheets, a cheaper option.

Task

Identify data, gather and process information to trace historical developments in the design and construction of ocean-going vessels with a focus on the metals used.

Hints and information that may assist in this investigation:

- To identify data sources, consider what is the most appropriate technology or strategy.
- Process the information by organising it in a chronological order. Note that the most important part of this investigation is to find the different metals used and the reasons for their uses.
- Some methods for the protection of iron from corrosion include:
 - iron covered in regular paint
 - iron covered in rust-preventing paints
 - iron covered in polymer-based paints
 - iron covered in vitreous enamel
 - tin plate (iron covered with a thin coating of tin)
 - galvanised iron (iron covered with a layer of zinc)
- Identify the ways in which a metal hull may be protected including:
 - corrosion resistant metals
 - development of surface alloys
 - new paints
- Stainless steels are favoured corrosion resistant metals because of a passive film of chromium (III) oxide on the surface that resists corrosion.
- Polymer paints protect against rust by forming a film over the surface of the steel that is impervious to oxygen and water. These paints also form a layer of a very insoluble ionic substance called pyroaurite. This ionic layer bonds strongly to the surface of the steel and well into the polymer layer. It prevents the movement of ions across the surface of the steel.

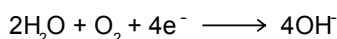
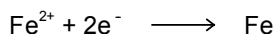
Activity 3: Rusting of iron in various aqueous solutions

An example of cathodic protection:

An iron ship may be protected from corrosion by attaching a block of magnesium to it. The magnesium block is easily replaced when it has nearly completely corroded. The magnesium is more active than the iron and corrodes first. The magnesium block and iron ship form an electrochemical cell and the magnesium is oxidised.



The electrons flow into the iron preventing the formation of Fe^{2+} ions. The electrons produced by the oxidation of the magnesium reduce any Fe^{2+} ions that form, back to Fe atoms. The site where reduction occurs is called the cathode and the method is called cathodic protection.

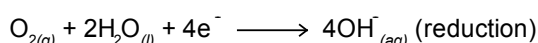
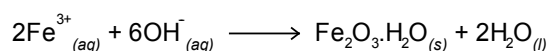
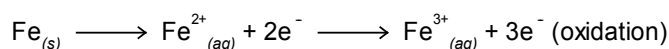


The rate of corrosion of the magnesium block is slow because it forms a coating of dull white magnesium hydroxide carbonate which protects the underlying magnesium.

In small boats sacrificial anodes are also used. Magnesium sacrificial anodes attached to outboard motors gradually dissolve as they are “sacrificed” instead of the metal parts of the outboard motor.

Corrosion is the process by which metals are converted to oxides or other compounds. This causes the metals to gradually deteriorate as illustrated by the rusting of iron and steel and the corrosion of aluminium fittings in salty ocean environments. Corrosion is an expensive problem in our society and quite extensive industries have developed that specialise in minimising the corrosion of metal structures.

In this experiment you will investigate the corrosion of iron and various factors that can accelerate or inhibit the rate of this process. In the rusting process iron is oxidised initially to iron(II) and then to iron(III). This combines with hydroxide ion, which is formed in the reduction of oxygen, to produce hydrated iron(III) oxide or brown rust ($\text{Fe}_2\text{O}_3 \cdot \text{H}_2\text{O}$). This can be represented by the equations



Equipment required

- 6 test tubes
- 10 clean iron nails
- 250 mL Beaker
- Bunsen, tripod and gauze mat
- Distilled water
- Vaseline
- Vaseline
- 10 mL Hydrochloric acid [HCl] 0.1 mol L⁻¹
- 10 mL Sodium hydroxide solution [NaOH] 0.1 mol L⁻¹
- 10 mL Sodium chloride solution [NaCl] 0.1 mol L⁻¹
- 10 mL Sodium dichromate solution [$\text{Na}_2\text{Cr}_2\text{O}_7$] 0.1 mol L⁻¹
- 2 mL Potassium hexacyanoferrate(III) solution [$\text{K}_3\text{Fe}(\text{CN})_6$] 0.1 mol L⁻¹

Procedure

1. Place a clean, bright nail into each of five test tubes.
2. Partly fill the test tubes with one of the following reagents so that the nail is just covered by the solution:
0.1 mol L⁻¹ HCl , 0.1 mol L⁻¹ NaOH , 0.1 mol L⁻¹ NaCl , 0.1 mol L⁻¹ $\text{Na}_2\text{Cr}_2\text{O}_7$, and distilled water.
3. Fill a sixth test tube with distilled water to a depth that will completely cover a nail. Boil the water vigorously for about 3 minutes to drive out dissolved air. While the water is still hot slide in a clean nail and add 1 mL of Vaseline to cover the surface of the water. The Vaseline melts, forming a layer between the air and the water.
4. Allow the nails to stand in the solutions for 24 hours.
5. After 24 hours observe the nails and solutions carefully and record any evidence of rusting.
6. To those solutions in which there is no evidence of rusting, add 2 drops of 0.1 mol L⁻¹ $\text{K}_3\text{Fe}(\text{CN})_6$ solution. The formation of a blue precipitate is indicative of the presence of iron(II) ions and, therefore, corrosion.

Safety note:

- Wear safety glasses when heating
- Handle all solutions with care avoiding contact with the skin and eyes, the acids and bases are corrosive

Results

1. Construct a table to summarise the results.
2. List the solutions showing evidence of corrosion. Indicate the relative amounts of corrosion in each case. Separately list those solutions in which there was no evidence of corrosion.
3. What substances must be present for rusting to occur? Describe any evidence from this experiment that supports this.
4. Explain the different amounts of rusting in the NaCl , HCl , and NaOH solutions when compared with the distilled water.

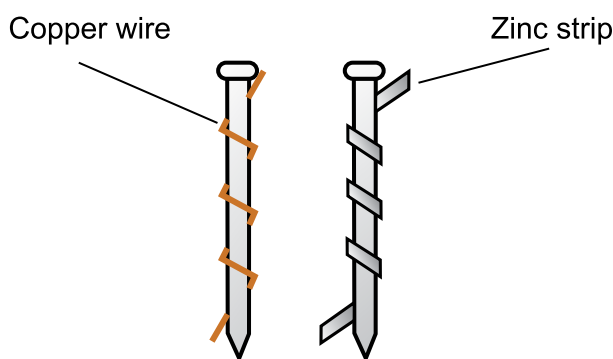
Activity 4: Rusting of iron in Contact with Other Metals

Equipment required

- Two test tubes and test tube rack
- 2 Clean iron nails
- 10 cm Copper wire
- 10 cm \times 2mm Zinc strip
- Distilled water

Procedure

1. In a test tube place a clean nail around which a piece of copper wire has been tightly wrapped
2. In a separate test tube place a clean nail around which a piece of zinc strip has been tightly wrapped.
3. Fill the test tubes with tap water and set them aside in a suitable location in the laboratory. Examine them at the end of the laboratory period and again after they have stood for 48 hours.



Processing of results, and questions

1. For the two metal couples studied, in which case was the iron corroded? Using the activity series of metals explain your results.
2. Explain why a nail can stand for many days on the shelf and not rust but will rust rapidly if placed in water.



Activity 5: Corrosion home practical

Corrosion of iron and its alloys:

The water of the oceans is a massive electrolyte that enables electrochemical reactions, including corrosion, to occur. Galvanic cells that lead to corrosion can form between different metals connected by sea water electrolyte.

Ocean going ships are constructed from iron or alloys of iron. The term steel is used to identify alloys of iron. Different steels have different properties because of their different composition. Rusting is the term applied to the corrosion of iron and steel and is an important chemical property of these substances.

Evidence about rusting is readily available from standard chemistry textbooks, encyclopaedias and from the World Wide Web. You need to identify the factors that affect the rusting of iron and be able to explain why they act in this way.

Some information you might find useful

- Both oxygen and water are necessary for rusting because together they can be reduced to form hydroxide ions.
- Salt water accelerates rusting because it is an electrolyte solution and its ions are able to move freely between the sites of the oxidation and reduction reactions.
- Impure iron rusts more quickly than pure iron because the impurities may act as the cathode. The O_2 is reduced on these impurities.
- The rusting of iron is accelerated when it is in contact with a less active metal such as copper or tin because less active metals serve as the cathode.
- Iron under mechanical stress rusts more quickly because individual Fe atoms are less strongly bonded together. This makes it easier for atoms to break away from the crystal lattice as Fe^{2+} ions.

Background information:

Iron will rust when it is in contact with both oxygen and water. If the water is acidic and if it contains dissolved salts (an electrolyte solution) then rusting is faster. Rusting of iron is also accelerated if the iron is impure, is in contact with a less active metal or is or has been under mechanical stress.

Plan your investigation:

You will need to research the chemical process of rusting.

Rusting of iron occurs when certain conditions are all present, such as carbon, salt, water and oxygen.

Design an experiment to investigate one factor that affects rusting. You could investigate the effects of controlling that factor in various ways.

Conduct your investigation:

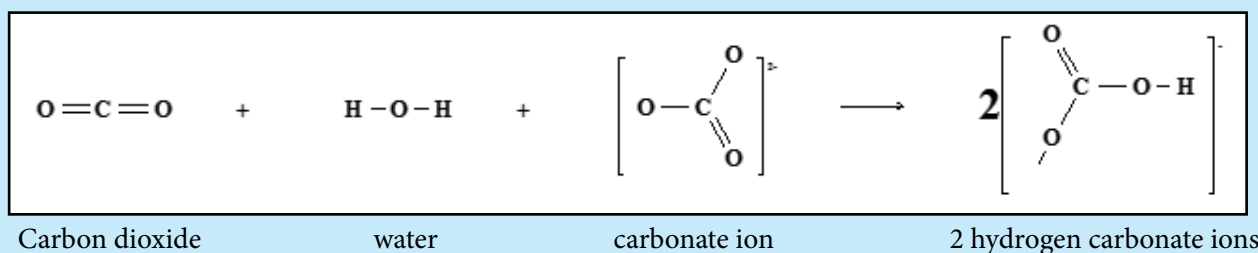
Carry out the experiment. The experiment will take many days or a week so you should start early.

Present your findings:

Write a report. Start the report at the planning stage. The report should have the following headings:

- Introduction:** Introduce your experiment and explain what corrosion is and how it happens.
- Aim/hypothesis:** Identify the variable (factor) you are trying to test and the results you expect.
- Risk Assessment:** Describe and explain all safety precautions that need to be followed.
- Procedure:** Show your planning. Clearly describe the steps you will follow and why you chose those steps. What you did, how long things were left, the containers you used, the conditions under which it was done. Remember someone else should be able to follow your procedure and come up with similar results.
- Results:** A table of results and we strongly suggest you include a sequence of photographs.
- Analysis of Results:** Analyse and discuss the results. Describe what went on and what the results prove. Evaluate the experiment; say what succeeded and what went wrong and how the experiment could be improved if you were to do it again.
- Conclusion:** The conclusion should address the aim / hypothesis.

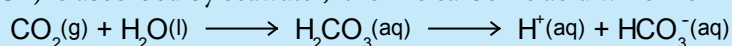
Ocean Acidification



The Chemistry

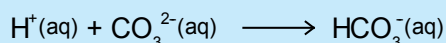
The ocean has many dissolved soluble salts; the major species being sodium ions and chloride ions. Substances we think of as 'insoluble' like calcium carbonate do dissolve in small quantities, 15 mg / litre is all that will dissolve and we call the ocean a saturated solution of calcium carbonate.

When carbon dioxide (CO₂) is absorbed by seawater, it forms carbonic acid which forms hydrogen ions (H⁺).



This formation of H⁺ (acid) decreases the pH of the water and is called "ocean acidification".

Calcium carbonate minerals are the building blocks for the skeletons and shells of many marine organisms. In the ocean, the seawater is supersaturated with respect to calcium carbonate. This means there are building blocks for calcifying organisms to build their skeletons and shells. However, continued ocean acidification is causing many parts of the ocean to become undersaturated with these minerals, which is likely to affect the ability of some organisms to produce and maintain their shells.



Since the beginning of the Industrial Revolution, the pH of surface ocean waters has fallen by 0.1 pH units. Since the pH scale is logarithmic, this change represents approximately a 30 percent increase in acidity. Future predictions indicate that the oceans will continue to absorb carbon dioxide and become even more acidic. Estimates of future carbon dioxide levels, based on business as usual emission scenarios, indicate that by the end of this century the surface waters of the ocean could be nearly 150 percent more acidic. This would result in a pH that the oceans haven't experienced for more than 20 million years.

The Biological Impacts

Ocean acidification is expected to impact ocean species to varying degrees. Photosynthetic algae and seagrasses may benefit from higher CO₂ conditions in the ocean, as they require CO₂ to live just like plants on land. On the other hand, studies have shown that a more acidic environment has a dramatic effect on some calcifying species, including oysters, clams, sea urchins, shallow water corals, deep sea corals, and calcareous plankton. When shelled organisms are at risk, the entire food web may also be at risk. Today, more than a billion people worldwide rely on food from the ocean as their primary source of protein.

Coral

Many marine organisms that produce calcium carbonate shells or skeletons are negatively impacted by increasing CO₂ levels and decreasing pH in seawater. For example, increasing ocean acidification has been shown to significantly reduce the ability of reef-building corals to produce their skeletons. Research results suggest that ocean acidification could severely impact the ability of coral reefs to recover from disturbance. Other research indicates that, by the end of this century, coral reefs may erode faster than they can be rebuilt. This could compromise the long-term viability of these ecosystems and perhaps impact the estimated one million species that depend on coral reef habitat.

Ocean Acidification: An Emerging Global Problem

Ocean acidification is an emerging global problem. Over the last decade, there has been much focus in the ocean science community on studying the potential impacts of ocean acidification. Since sustained efforts to monitor ocean acidification worldwide are only beginning, it is currently impossible to predict exactly how ocean acidification impacts will cascade throughout the marine food chain and affect the overall structure of marine ecosystems. With the pace of ocean acidification accelerating, scientists, resource managers, and policymakers recognize the urgent need to strengthen the science as a basis for sound decision-making and action.

Activity 6: The effect of ocean acidification on marine shell

The acidification of the ocean by carbon dioxide:

The ocean under normal conditions is a saturated solution of calcium carbonate. The calcium carbonate is available for use by corals and molluscs to build shell and coral reef.

This experiment models what can happen in our oceans as they become more acidic. In the experiment CO_2 in exhaled air represents the acidification of the oceans, while the saturated calcium hydroxide (limewater) solution represents the oceans. The extra CO_2 reacts with the 'available' calcium carbonate (milky solid precipitate) forming soluble calcium bicarbonate. The liquid becomes more acidic, less cloudy and orange in colour. The white solid, calcium carbonate becomes soluble colourless, calcium bicarbonate solution. In the oceans this means that less calcium carbonate is available to the corals and molluscs.

Aim:

To examine the acidification of the ocean by carbon dioxide and the impact on marine shell

Equipment required

- Universal Indicator™ Solution
- 5 mL Limewater - saturated $\text{Ca}(\text{OH})_2$ solution
- 10 mL Distilled water
- 100 mL Erlenmeyer flask
- Drinking straw

Procedure

1. Add 5 mL of limewater and 10 mL of distilled water to the flask.
2. Add 4 drops of Universal Indicator™ Solution. Note the colour and record your results (original properties) in the results table below. Use the Universal Indicator™ colour chart to record the pH of the solution.
3. Using the straw, gently blow through the limewater solution.
4. Note any colour changes seen and record when the precipitate begins to form. Note the colour at maximum milkiness.
5. Continue to blow air through the solution until the precipitate just disappears. Note the colour of the Universal Indicator™. This will take some time as exhaled air has very little carbon dioxide in it. That's why we can use it for mouth-to-mouth resuscitation.

Results: Complete the table

	Colour	pH
Original properties		
Milky precipitate first formed		
Maximum milkiness		
Milkiness started to disappear		
Most clear mixture		

Processing of results, and questions

1. Describe and explain the colour changes as you exhale through the straw into the limewater.
2. When carbon dioxide dissolves in water, it forms carbonic acid. Write the chemical equation for this process.
3. Write the name and chemical formula of the milky precipitate?
4. Explain with the use of chemical equations the formation of the milky precipitate.
5. Explain with the use of chemical equations the disappearance of the milky precipitate.
6. Compare the solubilities of calcium carbonate and calcium bicarbonate.
7. (a) Research and list causes of ocean acidification.
(b) What is the main cause of ocean acidification?
(c) Is this experiment a fair representation of the effect of ocean acidification on marine shell? Explain.



Early Colour



Finish colour