

Experiment 7

Desalination of Sea Water



The Task

The goal of this experiment is to investigate the nature and some properties of sea water.

Skills

At the end of the laboratory session you should be able to:

- use a balance to weigh samples,
- carry out a distillation,
- use a Bunsen burner,
- measure the conductivity of a liquid.

Other outcomes

You will develop an understanding of how water can be desalinated.

You will observe the effect of an electric current on water.

The Assessment

You will be assessed on your ability to set up the apparatus and perform a distillation correctly. See Skill 10.



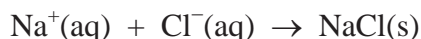
Introduction

Sea water is a very common substance, covering more than 70 % of the Earth's surface. The salinity of sea water is due to its containing a variety of dissolved ionic compounds. The main ions present are sodium ions (Na^+) and chloride ions (Cl^-), but there are also significant amounts of magnesium (Mg^{2+}), calcium (Ca^{2+}), potassium (K^+), sulfate (SO_4^{2-}) and hydrogencarbonate (HCO_3^-) ions. The level of salinity varies considerably around the world, approaching 4 % w/w (*i.e.* 4 g of total salts per 100 g of sea water) in sheltered seas with high evaporation rates and low rainfall (*e.g.* the Red and Mediterranean Seas) and falling below 1 % w/w in regions of high fresh water input (*e.g.* the Baltic Sea, polar seas and near the mouths of the largest rivers). The global average is generally considered to be 3.5 % w/w and that is typical of the seas around Australia. Unfortunately for Australia, the quotation from the Rime of the Ancient Mariner (Samuel Taylor Coleridge, 1798) is often appropriate: “*Water, water, everywhere, nor any drop to drink.*” Australia does not have large reserves of drinkable freshwater and is constantly threatened by drought. Therefore, the desalination of its plentiful supplies of sea water appears to be necessary to secure a stable water supply for the future.

Hydration, Evaporation and Crystallisation

Ionic compounds in solution are present as hydrated ions, freely moving around in the solvent. The cations, *e.g.* Na^+ , are surrounded by water molecules which are oriented so that the slightly negatively charged O atoms are closer to the ion than the slightly positively charged H atoms. The anions, *e.g.* Cl^- , are also hydrated, but the water molecules are oriented in the opposite way so that the H atoms are closer to the negative ion. To emphasise that the ions are hydrated, it is common to write their formulae with an (aq) attached thus: $\text{Na}^+(\text{aq})$, $\text{Cl}^-(\text{aq})$. Hydration of ions arises from electrostatic attraction between the charge on the ion and the oppositely charged end of the polar water molecules.

If the water is removed from a solution of an ionic compound, *e.g.* by boiling it off, the concentration of ions increases until the solution becomes saturated. Further removal of water causes the ions to combine to form crystals of the solid compound. The process is represented by the following equation.



If the water is removed rapidly from the solution, the resulting crystals are usually small, like normal table salt. Alternatively, slow loss of water from a solution, *e.g.* by allowing it to evaporate, can lead to larger crystals of the ionic compound, such as those seen in the “rock salt” used in salt grinders.

Distillation

Distillation is a process of heating a liquid until it boils, then collecting the vapour and condensing it back into a liquid. It is used most effectively to purify mixtures whose components have very different boiling points. Water boils at 100 °C and sodium chloride, being a typical inorganic salt with high melting and boiling points, boils at over 1400 °C. Most ionic solids will not even melt at temperatures available from a Bunsen flame, so distillation is an excellent method for separating water from the dissolved salts in sea water. As the water is distilled off from the sea water and the salt concentration in the remaining solution increases, the ions in solution combine to form solid compounds which remain in the flask.

Conductivity and Electrolysis

A substance can conduct electricity if it contains charged particles that can move freely. In metals, these particles are the outer (valence) electrons, which are able to move freely from atom to atom when a voltage is applied. This process, called metallic conduction, does not alter the metal chemically.

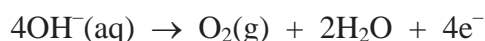
A substance that consists of positively and negatively charged ions can also conduct electricity, but only if those ions can move freely. Solid ionic compounds are made up of ions in a crystal lattice and do not conduct electricity because the ions, which are firmly anchored in the lattice, cannot move in response to an applied voltage. However, conduction does occur when an ionic compound is molten or when it is dissolved in a suitable solvent (*e.g.* water). The ions move by electrostatic attraction to the electrode of charge opposite to that on the ion. At the electrodes, the ions undergo chemical reactions, known as *reduction* and *oxidation*, in which electrons are transferred. Conduction of a current through an ionic liquid or solution therefore results in changes in the compound(s) involved. This process is called *electrolysis*.

Water consists of H₂O molecules. A very small proportion of these molecules (about 1 in 56 million) splits up into H⁺ and OH⁻ ions. Although in this experiment you will not be able to detect the electrical conductivity of pure water, very sensitive conductivity equipment can reveal the presence of ions.

By definition, the *cathode* is the electrode at which the reduction reaction (a gain of electrons) occurs. Here, that is the negative electrode, where the H⁺ ions gain electrons to form hydrogen gas:



By definition, the *anode* is the electrode at which the oxidation reaction (a loss of electrons) occurs. Here, that is the positive electrode, where the OH⁻ ions lose electrons to form oxygen and water:



Electrons produced at the anode are taken up by the positive terminal. At the same time, electrons supplied from the negative terminal react at the cathode. As the H⁺ and OH⁻ ions are removed by these oxidation and reduction reactions, they are constantly replaced by more water molecules dissociating into H⁺ and OH⁻ ions. This is an example of an electrolytic cell.

Safety

Chemical Hazard Identification

5 M sulfuric acid - corrosive, avoid contact with your skin.

Salt water - non-hazardous.

Risk Assessment and Control

Moderate Risk.

Magnetic stirrer hotplates are a potential burn hazard.

The production of steam and hot equipment creates a risk of burning and scalding.

Waste Disposal

All of the solutions can be disposed of down the sink. Solid waste should be placed in the bin.

Experimental

This experiment is to be carried out in pairs.

Note: Rather than use actual sea water, you will carry out the experiments on solutions of salt water, specially prepared for the experiment.

Part A Distillation

(A1) Set up the distillation apparatus as described in Skill 8. Have a demonstrator check it before you start the distillation.

(A2) Pour 75 mL of salt water or deionised water (your Demonstrator will tell you which) into a 100 mL round-bottomed flask and add 3 anti-bumping granules to encourage smooth boiling.

(A3) Place a clean dry 100 mL conical flask under the end of the delivery tube and begin heating the flask with a heating mantle. Heat the flask until you have collected about 60 mL of water at a rate of about 2 drops per second. Record the temperature of the boiling water vapour in your logbook.

(A4) Turn off the heating mantle and allow the flask to cool.

(A5) Save the distillate (*i.e.* the liquid that has distilled over) in the 100 mL conical flask for a later part of the experiment. Seal it with plastic wrap and label the flask with its contents and your names.

For your logbook:

What was the temperature of the boiling water vapour? If it was not 100 °C, can you explain why?

Part B Determination of the salt concentration in salt water

(B1) Weigh a clean 250 mL beaker containing another 3 fresh anti-bumping granules and covered with an 8 cm watch glass. Record the total mass of the beaker, granules and watch glass in your logbook.

(B2) Pour approximately 30 mL of salt water into the beaker. Reweigh the beaker with the granules, watch glass and now including the water. Record the total mass in your logbook. Calculate the mass of salt water and record it in your logbook.

(B3) Put the beaker on a stirrer hotplate and heat to boiling point. Maintain the correct setting to promote steady boiling. When the volume has been reduced to about 70 %, place the watch glass partially over the beaker to prevent splashing and continue to heat the beaker until all the water has just evaporated.

Caution: Do not continue to heat the beaker after the last trace of water has evaporated or it will crack.

(B4) Carefully remove the beaker using metal tongs and place on a fibreglass mat to cool.

(B5) Allow the beaker to cool and then weigh the beaker and its contents along with the watch glass. Record the mass in your logbook.

For your logbook:

Calculate the mass of the salt that has crystallised.

Calculate the weight percentage (% w/w) of salt in the salt water.

How does your value compare with what you know about typical salt contents of sea water (see Introduction)?

Part C Density of salt water

(C1) Measure the densities of deionised water and salt water using the hydrometers in the fume hood.

For your logbook:

Do the densities you measured from the hydrometers agree with the weight percentages you calculated in Part B? If not, why not?

Part D Conductivity of distilled water compared with salt solutions

To investigate the conductivity of a liquid, a voltage is applied across it. A small light globe included in the circuit glows if a current flows, and the brightness of the globe gives a qualitative assessment of the degree of conductivity of the liquid. If there is no conduction, the globe doesn't glow.

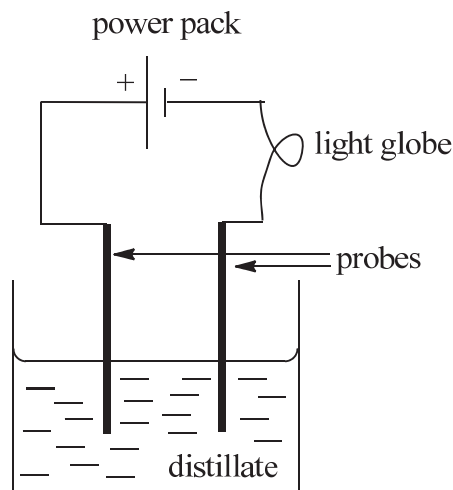


Figure 1

(D1) Place about 50 mL of the distillate you obtained from the distillation of sea water in Part A in a clean, dry beaker. Set up a circuit as shown in Figure 1, with the probe electrodes placed in the water in the beaker. Switch the current on and check for an indication of conduction. Record your observations in your logbook. Don't discard the distillate. You'll need it again in Part E of the experiment.

(D2) Repeat the same procedure using salt water instead of distilled water. Again record your observations in your logbook, noting in particular any differences between the behaviour you observed for the distilled water.

(D3) Repeat the same procedure again with a solution made by dissolving a spatulaful of sodium chloride in 50 mL of deionised water. Record your observations in your logbook.

(D4) Repeat the procedure once more on a solution made by dissolving a spatulaful of sucrose (sugar) in 50 mL of deionised water. Record your observations in your logbook.

For your logbook:

Rank the solutions in the order of their conductivity.

Can you propose a hypothesis to explain the order that you observed?

Part E Electrolysis of distilled water

- (E1) Mix 50 mL of the distillate you obtained from Part A with 5 mL of 5 M H_2SO_4 .
- (E2) Place the solution in the electrolysis unit (Figure 2). Bend the rubber tubing over on the top of each arm of the electrolysis unit and clamp it to avoid gas escaping. Have the apparatus checked by your demonstrator.

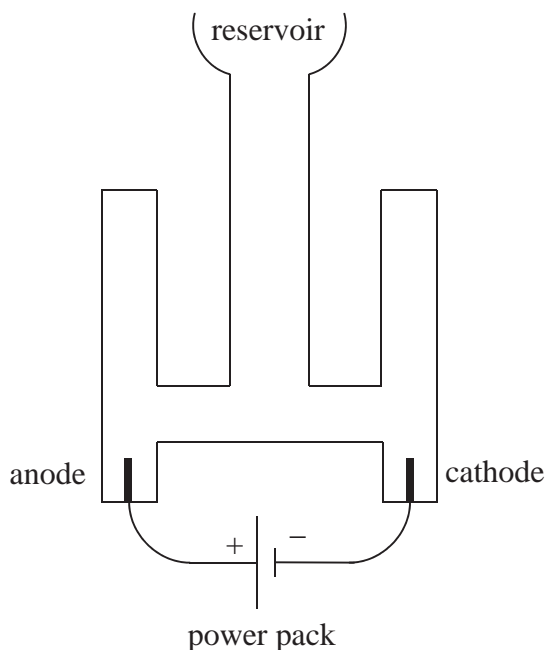


Figure 2: Set up of the electrolysis unit (Hoffman voltameter)

- (E3) Connect the power to the electrodes. When you have collected at least 2 cm of gas in each arm, disconnect the power from the electrodes.
- (E4) Estimate the relative volumes of gas at each electrode and record this ratio in your logbook.
- (E5) Test the gas at the cathode by placing an inverted test-tube over the end of the cathode outlet tube and releasing the gas into the test-tube. Light a long thin candle and then hold it under the mouth of the still-inverted test-tube. Describe your observation in your logbook.
- (E6) Reconnect the power and continue the electrolysis until the gas from the anode fills the collection chamber. Disconnect the power from the electrodes.
- (E7) Light a long thin candle with a match and then blow it out so that there are still some embers glowing. Immediately hold it near the rubber tube at the top of the anode chamber. Release the gas from the chamber (without allowing the solution to flow out) so that it passes over the embers. Describe your observations in your logbook.

For your logbook:

Which gases do you think were produced at the anode and the cathode, respectively?

Can you explain the relative volumes of gases initially produced at the anode and cathode in (E4)?

Write a balanced equation for the electrolysis of water.

Suggest why H_2SO_4 was added initially to the solution.

Group Discussion

Was there any difference between the boiling points of salt water and pure water? If so, can you explain the difference?

In principle, sea water could be desalinated by distillation. However, Sydney's new desalination plant at Kurnell doesn't use this technique. It actually uses the technique of "reverse osmosis". Does anyone know what osmosis is? Can anyone suggest what *reverse osmosis* might be?

Would there be any difference in the energy change between sea water and fresh water when desalination is carried out via distillation or reverse osmosis? If not, why do you think desalination plants prefer to use the technique of reverse osmosis over distillation?