

CARBON CAPTURED CHALK

Tony Butterfield, based off of a module developed by Alissa Park. Developed by The NSF Research Coordination Network on CCUS (<http://www.aiche.org/ccus-network>). | October 14, 2013



SUMMARY:

This module demonstrates several important chemical engineering concepts. Two solutions of soluble salts (Na_2CO_3 and CaCl_2) are mixed and undergo a double substitution reaction. The resulting product is table salt (NaCl) and chalk (calcium carbonate). Concepts of solubility, enthalpy of solutions, reaction kinetics are considered. A very similar process could be used by chemical engineers to help relieve the burden of carbon dioxide on our atmosphere by trapping our emissions in benign products, such as chalk.

IMPORTANT!!! Under no circumstances should an unsupervised minor perform the procedures described herein. All the following described experiments and methods should be supervised by an adult who is completely familiar with and takes full responsibility for all possible hazards.

GENERAL INFORMATION:

Main Curriculum Tie: Chemistry

Additional Curriculum Ties: Separations, Environmental Science

Career Connections: With growing concern over global climate change, many industries and governmental bodies have become interested in development of technologies that may be used to keep greenhouse gases out of our atmosphere. Chemical, environmental, and civil engineers play a critical role solving climate problems, and this module gives a hands-on demonstration of one such technology.

Mean Time Frame: 30 min

Group Size: Teams of 2-3, 50 students max

Student Prior Knowledge: Basic Chemistry

ESSENTIAL QUESTIONS

- What is a double replacement reaction?
- What is solubility?
- How can we remove carbon dioxide from our atmosphere?

BIBLIOGRAPHY

Alissa Park, OSCAR (Ohio State Carbonation and Ash Reactivation) Process, teaching module.

MATERIALS & METHODS

Materials

- Sodium Carbonate (aka Na_2CO_3 , soda ash, washing soda). Can be purchased from chemical supplier or even at [amazon.com](https://www.amazon.com).
- Calcium Chloride (aka CaCl_2). Can again be purchased from a chemical supplier, or more easily as ice melt ([amazon.com](https://www.amazon.com)).
- Water.
- Scale, some means to measure weight. We bought [a cheap one](https://www.amazon.com) from amazon.com and it works great (This module seemingly brought to you by amazon.com...).
- Weighing boats or weighing paper to keep your scale from getting dirty.

- Spatula, to finely measure out solids on the scale.
- 50 ml disposable centrifuge tube, or any container that can hold the volume of liquid needed. The centrifuge tubes are good because they can be sealed and shaken to dissolve solids. A stir plate could also be used with, say, a beaker.
- Filter paper (we're using #4 filter paper, but coffee filters can work passably). If you want to filter quickly a vacuum filter could be added but gravity filtration works here too.
- Safety glasses and gloves. This is a pretty safe activity, but it's always good to use gloves and glasses in a lab environment.

Methods

1. Place a weighing boat on the scale and tare it out.
2. Measure out 8g of Na_2CO_3 and dump it into an empty 50 ml centrifuge tube. You may wish to begin this module with the [Carbon Capture in Water](#) module to illustrate how CO_2 might get from the atmosphere into a compound such as sodium carbonate.
3. In the same manner, measure out 8 g of CaCl_2 . Alternatively, students can calculate exactly how many grams of CaCl_2 needed to react one-to-one with the Na_2CO_3 , and measure that out, but it will be near 8 g anyway.
4. Dump the CaCl_2 into an empty 50 ml centrifuge tube.
5. With a third 50 ml centrifuge tube or a graduated cylinder, measure out 20 ml of water and pour that water in with the Na_2CO_3 .
6. The solution will heat up as the Na_2CO_3 dissolves. Discuss this phenomena with the students.
7. To completely dissolve the salt put the lid on the tube and shake the solution, periodically stopping to open the lid to release any pressure that may build up due to the heat.
8. Measure out the same amount of water and mix it with the CaCl_2 in the other centrifuge tube. Mix in the same manner as in Step 7.
9. You should now have two tubes with two different solutions in them.
10. Pour the CaCl_2 solution into the centrifuge tube containing the Na_2CO_3 solution. If you pour slowly students can see the difference in solution density and can witness their interaction at the interface between the two.
11. Put the lid on the tube and mix the two solutions together by shaking. You will notice the solution go opaque as it initially forms a gel. After about 3 min of shaking, the solution should become far less viscous and look a lot like milk as the chalk forms minute particles.
12. Separate the water from the chalk by folding filter paper into a cone over an empty 50 ml centrifuge tube and pouring your solution into it. The water should be collected in the tube and the chalk will remain in the filter paper. You may also wish to use vacuum filtration or a centrifuge if you have it.

13. Finally, place the chalk out to dry and by the next day you should have a dried product.

BACKGROUND FOR TEACHERS

The theory behind this module is based on several relatively simple concepts.

Enthalpy of Solution

In this module twice we start with a dry salt and end up with that salt dissolved into solution. Consider what happens to the atoms in CaCl_2 during this process. Before they can enter the liquid they have to disassociate from the CaCl_2 surrounding them. They must break bonds in their crystalline structure, which may be seen as pits of potential energy that they have to escape; this takes energy. We call the energy needed the *lattice dissociation enthalpy*.

To get into solution an atom from CaCl_2 must next become surrounded by solvent molecules, which is water in our case. The energy required to do this is called the *enthalpy of hydration*, and it is always negative, meaning that the potential energy of the system is always lower with the salt ions surrounded by water than if they are disassociated in a vacuum. This also means that energy is released in this process.

So we have a process that takes energy (breaking the salt's lattice), and a process that releases energy (surrounding ions by water). The difference between the two will result in either an energy release or uptake, a temperature increase or decrease. If the enthalpy of hydration is more negative than the lattice dissociation enthalpy is position, then the total enthalpy change will be negative and energy will be released into the system, resulting in a temperature increase. This is the case we have with both our salts, CaCl_2 and Na_2CO_3 (though this is not the case for all solutes). As the students dissolve both of them, there will be a notable increase in temperature in the solution.

Double Replacement Reaction & Solubility

While in solution, CaCl_2 dissociates into two ions of chlorine and one ion of calcium, while Na_2CO_3 dissociates into two ions of sodium and one carbonate ion. These ions may interact with each other, and once a calcium ion meets up with a carbonate ion they form calcium carbonate (chalk) and sodium chloride, table salt.



This reaction strongly favors the forward direction. Chalk is not very soluble in water (it's solubility is about 13 mg/L) as one could easily discover with a piece of chalk and a glass of water. It takes less energy for calcium chloride to interact with itself than to interact with molecules of water in solution. As such, the concentration of chalk in aqueous form is continually held low as it precipitates out of solution to form a solid which will not participate in the reaction.

Because a product in the reaction is continually being removed from solution, and the rate of reaction in the reverse direction depends on the concentration of calcium carbonate, the reaction shown above is forced to the right.

VIDEOS

- YouTube Videos:
 - <http://www.youtube.com/user/UnvUtahChenEng>

INTENDED LEARNING OUTCOMES

- Students will understand the difference between soluble and insoluble.
- Students will be able to define heats of solutions.
- Students will be able to identify a means to sequester carbon dioxide.

INSTRUCTIONAL PROCEDURES

1. Discuss the importance of carbon sequestration technology.
2. Explain how CO_2 may be converted into carbonic acid and then sodium carbonate.
3. Split the class into teams of two or three.
4. Assign out tasks as described in the Methods section of this module within each group.
5. Ask students to pay attention to the temperature of each solution as they dissolve their salts.
6. Ask students to pay attention to and make observations of the viscosity of their mixture of the two solutions.
7. Finish the methods and, as the chalk filters, ask students to perform the assessment questions in this module.

OPTIONAL ACTIVITIES & EXTENSIONS

- Couple this module with our [Carbon Capture in Water](#) module to illustrate how carbon dioxide can be removed from the atmosphere and put into substances such as the sodium carbonate reactant used in this module.
- Have students put a drop of each reactant solutions on a microscope slide and have them observe the microscopic formation of calcium carbonate crystals.
- Allow students to dye their solution and then press the resulting chalk into a solid piece of chalk as it dries.

ASSESSMENT PLAN

Possible questions to be used for assessment:

Q1. The solubility of CaCl_2 in water at room temperature is 740 g/L, and the solubility of Na_2CO_3 in water at room temperature is 215 g/L. How many grams of each could go into separate 20ml volumes of water?

- A1.

For CaCl_2 :

$$740 \text{ g/L} * 0.02 \text{ L} = \mathbf{14.8 \text{ g}}$$

For Na_2CO_3 :

$$215 \text{ g/L} * 0.02 \text{ L} = \mathbf{4.3 \text{ g}}$$

(The students should notice this is less than the 8g they dissolved in 20ml of water in this module)

Q2. What mass in grams of Na_2CO_3 will we need to react with 10g of CaCl_2 ?

- A2.

$$\text{MW of } \text{CaCl}_2 = 40 + 2*35.5 = 111 \text{ g/mol}$$

$$\text{MW of } \text{Na}_2\text{CO}_3 = 2*23 + 12 + 3*16 = 106 \text{ g/mol}$$

Because the molecular weights are so close we should expect the needed mass to be near 10g...

Moles of Na_2CO_3 needed =

$$10\text{g} / 111\text{g/mol} = 0.09 \text{ mol}$$

Grams of Na_2CO_3 needed =

$$0.09\text{g} * 106\text{g/mol} = \mathbf{9.55 \text{ g}}$$

Q3. In a 1 L volume, we react 1 mol/L of silver nitrate (AgNO_3) with equimolar proportions of sodium chloride (NaCl) in a double substitution reaction. What should be the products and what mass of products will be created if the reaction goes to completion? Do the products remain in solution or precipitate?

- A3.



The products will be silver chloride (AgCl) and sodium nitrate (NaNO_3). By the balanced reaction it's clear that we should end up with 1 mol/L AgCl and 1 mol/L NaNO_3 . As the molecular weight of AgCl is 143 g/mol, we would end up with 143 g/L of AgCl , which would be **143 g** in 1 L. Similarly, we would end up with **85 g** of NaNO_3 .

If the students look up the solubility of silver chloride in water at room temperature they'll find it to be very low (about 5 mg/L), while the solubility of sodium nitrate is at about 900 g/L. **So all but 5 mg of the AgCl will precipitate out, while all the sodium nitrate will remain dissolved.**

Q4. We wish to remove gaseous carbon dioxide from the exhaust of a coal power plant and convert it into solid chalk. We will first react CO_2 with water containing sodium hydroxide. We should form

carbonic acid which will then quickly react with the sodium hydroxide to form sodium carbonate. From there we will add calcium chloride to finally produce calcium carbonate.

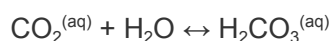
List these processes in order, following the carbon, and balance the three primary reactions involved mentioned above. Give the overall reaction for this process. Finally, calculate the number of grams of reactants we will need, and the grams of products we'll create for each gram of carbon dioxide removed from the exhaust.

- A4.

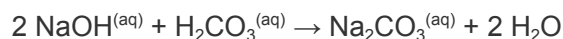
First the carbon dioxide must leave the gas and dissolve into the water



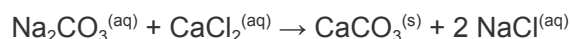
Next the dissolved carbon dioxide will react with water to form carbonic acid:



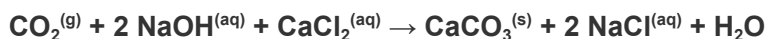
Very quickly we will remove the carbonic acid by its reaction with sodium hydroxide:



Finally we react the sodium carbonate with calcium chloride:



Adding these all up, the overall reaction would be



To get the mass of reactants needed and products produced we simply need to use the overall reaction and the ratios of molecular weights multiplied by their stoichiometric coefficients. For example, it is clear that for each molecule of carbon dioxide (MW = 44 g/mol) we will create a molecule of calcium carbonate (MW = 100 g/mol). So for each gram of carbon dioxide we will make:

$$(1 \text{ g CO}_2^{(g)}) ((1*100)/(1*44)) = \mathbf{2.27 \text{ g CaCO}_3^{(s)}}$$

For the other compounds:

Formula	MW (g/mol)	Stoichiometric Coefficient	Mass (g/g CO ₂)
CO ₂ ^(g)	44	1	-1
NaOH ^(aq)	40	2	-1.82
CaCl ₂ ^(aq)	111	1	-2.52
CaCO ₃ ^(s)	100	1	2.27

NaCl ^(aq)	58.4	2	2.65
H ₂ O	18	1	0.41

Q5. What does it take to offset your personal carbon footprint? Go to the EPA's carbon footprint calculator (<http://www.epa.gov/climatechange/ghgemissions/ind-calculator.html>) or some [other carbon footprint calculator](#) if you are outside of the United States. Calculate the carbon footprint for your household. If you don't know the answers to some of the questions, for the EPA calculator, you can use the national average given below each text box.

We wish to remove your personal carbon dioxide from the atmosphere by reacting it with sodium hydroxide and calcium chloride to form calcium carbonate in a process similar to the one described in the previous question. How many pounds of reactants would we need to purchase and how much of each product would we create (in lb)? Finally, design a rough draft of a chemical plant that could handle this task and indicate the flow rates of materials into this plant as lb per day.

- A5.

Each student's answer will be different, due to their different carbon footprint, so we will do the calculation for the average American. Let's assume the EPA site is right, and the average American household produces 83,000 lb of carbon dioxide each year.

The first part of this solution should result in the table produced in the previous assessment question. By simply multiplying the ratios of masses in the last column by the total mass of carbon dioxide produced we can get the masses of the reactants needed and products. Note that (g/g CO₂) is the same as (lb/lb CO₂).

Formula	MW (g/mol)	Stoichiometric Coefficient	Mass Ratio (lbs/lbs CO ₂)	Mass (lb)
CO ₂ ^(g)	44	1	-1	-83,000
NaOH ^(aq)	40	2	-1.82	-151,000
CaCl ₂ ^(aq)	111	1	-2.52	-209,000
CaCO ₃ ^(s)	100	1	2.27	188,000
NaCl ^(aq)	58.4	2	2.65	220,000
H ₂ O	18	1	0.41	34,000

So to remove the carbon dioxide produced by the average American, each year we'd need 76 tons (1 ton = 2000 lb) of sodium hydroxide, and 105 tons of calcium chloride. That could be pricey... But we'd also produce 94 tons of calcium carbonate which can be put to use, along with the 110 tons of table salt...

Clearly the process needed to do this would need to be quite large. Each day we'd need to process 227 lbs of CO₂, 573 lbs of CaCl₂ and so on.

The plant designs students come up with should be fairly simple but should indicate such significant flow rates. The process should have a mixing unit where water is mixed with NaOH. That basic water should go to some sort of gas stripping unit, where air is bubbled up through falling liquid and CO_2 is removed. The liquid solution should then go to a reactor where an appropriate amount of CaCl_2 is added. From there solid chalk will form and the slurry should be sent to some sort of separation unit, like a filter or settling tank. Finally the chalk should be dried in some manner. Advanced designs could have recycled streams to improve conversion and heat exchangers to take advantage of the heat generated. More advanced students might also want to consider the purity of the products that are produced and how well the separation units operate.