

## REACTING AN ALKA SELTZER™ TABLET

<b>TOPIC</b>	Measuring reaction rates	<b>DEMO #</b> 12.I.1
<b>REFERENCE</b>	commonly known	
<b>EQUIPMENT</b>	top loading automatic balance 250 mL beaker	
<b>CHEMICALS</b>	Alka Seltzer™ tablet room temperature	
<b>PROCEDURE</b>	Place a 250 mL beaker half-filled with room temperature water on the balance and “zero” the balance. Make a table with headings “Mass (g)” and “Time (s)” on the blackboard or overhead. Place the Alka Seltzer tablet on the balance and record the mass of the dry tablet in the 1st column and put “0” in the time column. Have one student carefully (so as to avoid splashing) “plop” the tablet into the beaker on a signal from a 2nd student who is acting as a timer and a recorder. Record the mass every 30 s until the student observing the progress of the reaction declares that the tablet has disappeared. Some gas will continue to rise to the surface after the reaction is ended.	

Plotting the data on a graph shows a more-or-less straight line with a negative slope. Students can be asked to calculate a value for the “mass lost per second”, including the units. Discussion shows that the slope of the graph is the rate of the reaction.

Students may ask about the fluffy white solid which remains in the beaker after the reaction. Read the label on the package to show them that the tablet contains “heat treated sodium bicarbonate” (dried baking soda), acetylsalicylic acid (“aspirin”) and citric acid. The white solid residue is undissolved aspirin and the purpose of the baking soda and salicylic acid is to produce carbon dioxide gas that convinces the user, after a satisfying burp, that the tablet is doing them some tangible good.

## AN EGG-SPLOSIVE DEMONSTRATION

<b>TOPIC</b>	Effect of concentration on reaction rate	<b>DEMO #</b> 12.I.2
<b>REFERENCE</b>	Twenty Demonstrations Guaranteed to Knock Your Socks Off! , p. 30	
<b>EQUIPMENT</b>	1 inch piece of electrician's tape egg, blown out (any size): see procedure 2 L plastic bottle safety shield plastic tube and plastic pipette (to deliver hydrogen into the tiny hole in the egg)	
<b>CHEMICALS</b>	source of hydrogen gas	
<b>PROCEDURE</b>	<p>Make a 2–3 mm hole at the top of the raw egg and a slightly larger (3–4 mm) hole at the bottom. Insert a wire and stir up the yolk. Blow into the small hole to force out the egg's contents, rinse several times and allow to dry overnight or at 100°C for 40 minutes.</p> <p>Cover the top hole with a piece of tape and flush out the air and fill with hydrogen using a delivery tube that extends almost to the top of the egg.</p> <p>Keeping a safety shield between the egg and the audience, stand the egg on the open top of the pop bottle, with the top of the egg up. Warn the students to “cup” their ears. To initiate the reaction, remove the tape and hold a burning match to the top hole and quickly stand back. There will be an initial, almost inaudible “pop” and then nothing appears to happen. Then after about 10–15 seconds ... BOOM and the egg is gone.</p>	

## THE EFFECT OF “THE NATURE OF THE REACTANTS”

<b>TOPIC</b>	Nature of reactants	<b>DEMO #</b> 12.I.3
<b>REFERENCE</b>	commonly known	
<b>EQUIPMENT</b>	2 – 250 mL beakers	
<b>CHEMICALS</b>	2 – clean strips of copper (about 2 cm x 15 cm each) bottle of concentrated nitric acid bottle of concentrated hydrochloric acid	
<b>PROCEDURE</b>	IN A FUME HOOD: pour about 25 mL of nitric acid into one beaker and 25 mL of hydrochloric acid into the second beaker. Simultaneously: place one copper strip into each beaker. Show that it is nitric acid’s “nature” to react with copper while hydrochloric acid does not.	

You may find it useful to comment on the fact that many war movies in the past have featured retreating soldiers going under bridges that had to be destroyed by pouring acid on the steel girders. The bridge was supposed to collapse just as the pursuing armies crossed the bridge. Point out that no acid can do more than corrode the small amount of the surface metal in contact with the few millilitres that adhere to the surface: hydrochloric acid reacts very slowly with iron and nitric acid, when concentrated, makes the iron “passive” and unreactive. Sorry, Hollywood!

## LYCOPODIUM EXPLODER

**TOPIC** Effect of surface area on reaction rate 1 **DEMO #** 12.I.4

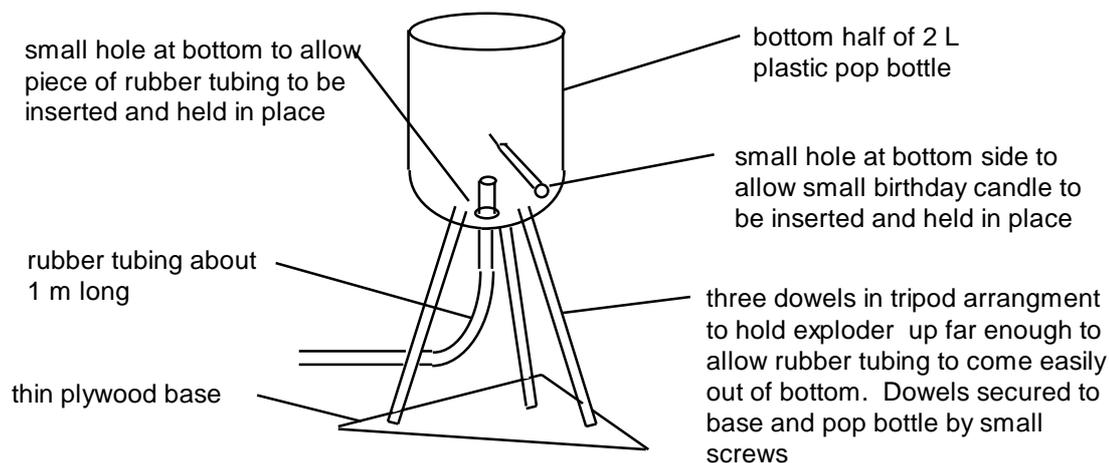
**REFERENCE** Chemistry Demonstration Aids That You Can Build: p. 61

**EQUIPMENT** lycopodium exploder (see reference or below)  
small birthday candle  
match  
rubber tubing: 3/8" OD, 5 feet  
wide stem Beral pipette  
sponge (for clean up)

**CHEMICALS** Bottle of lycopodium powder

**PROCEDURE** Insert the tubing into the bottom of the plastic bottle. Arrange the candle so the wick is in the centre of the bottle. Use the pipette to place 3–5 "loads" of lycopodium into the bottle end of the tubing. Make sure the end of the tubing is aimed at the end of the candle and light the candle. Turn off the room lights: take a large breath and blow into the other end of the tubing. A huge fireball erupts above the bottle.

Note: The exploder is simple to construct, as shown below.

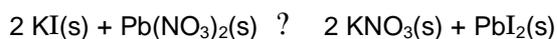


## STARCH EXPLOSION

<b>TOPIC</b>	Effect of surface area on reaction rate 2	<b>DEMO #</b> 12.I.5
<b>REFERENCE</b>	A Demonstration–A–Day...For High School Chemistry, p. 20	
<b>EQUIPMENT</b>	balloon funnel candle metre stick match	
<b>CHEMICALS</b>	cornstarch oxygen tank	
<b>PROCEDURE</b>	Use a funnel to put about a teaspoonful of corn starch into a balloon. Fill the balloon with oxygen, tie off the balloon and shake it vigorously. Avoid any possibility of static sparks!  Tape a candle to the end of a metre stick and light the candle. Move the candle under the balloon after the balloon is freshly shaken. BANG!	

## METATHESIS REACTION BETWEEN TWO SOLIDS

<b>TOPIC</b>	Reaction rate between solids	<b>DEMO #</b> 12.I.6
<b>REFERENCE</b>	Chemical Demonstrations: A Sourcebook for Teachers, Volume 1, p. 115	
<b>EQUIPMENT</b>	stirring rod 100 mL beaker	
<b>CHEMICALS</b>	about 2–5 g of potassium iodide crystals about 2–5 g of lead (II) nitrate crystals	
<b>PROCEDURE</b>	Place equal amounts of potassium iodide and lead (II) nitrate crystals in a 100 mL beaker. Stir the mixture with a stirring rod. After a short while, the mixture will take on a yellow colour.	



To show that this yellow colour is a new product, add water and notice the immediate formation of a bright yellow product. This also shows that ion mobility is much greater in liquids than in solids.

## DECOMPOSITION OF HYDROGEN PEROXIDE BY $\text{MnO}_2$

<b>TOPIC</b>	Catalysis 1	<b>DEMO #</b> 12.I.7
<b>REFERENCE</b>	commonly known	
<b>EQUIPMENT</b>	250 mL erlenmeyer flask spatula sponge (for clean up)	
<b>CHEMICALS</b>	100 mL 30% hydrogen peroxide powdered manganese dioxide	
<b>PROCEDURE</b>	Place 100 mL of hydrogen peroxide in the flask. Keep students at least 10 feet away. Quickly drop in about 1 g of $\text{MnO}_2$ . A rapid rush of oxygen is given off, then the reaction seems to slow but as sufficient heat accumulates the reaction starts up again, spewing out gas, steam and water. Point out to students that the mass of $\text{MnO}_2$ at the end is identical to the original mass used. Ask what function is served by the $\text{MnO}_2$ . (Catalyst)	

## HYDROGEN GAS PRODUCTION — A CATALYZING EXPERIENCE

<b>TOPIC</b>	Catalysis 2	<b>DEMO #</b> 12.I.8
<b>REFERENCE</b>	A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 227	
<b>EQUIPMENT</b>	2 – 500 mL erlenmeyer flasks 2 – balloons (capable of being stretched over the mouth of the flasks) 2 – 250 mL graduated cylinders	
<b>CHEMICALS</b>	mossy zinc 500 mL of 2 M sulphuric acid, $\text{H}_2\text{SO}_4$ three vials, each containing 0.2 g of copper (II) sulphate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$	
<b>PROCEDURE</b>	Place 150 mL of sulphuric acid into each flask. To one flask add 0.2 g of copper sulphate and swirl to dissolve. Put one large piece of zinc into each flask and quickly put a balloon over the top of each flask. The balloon on the flask containing the copper catalyst quickly swells, the other much more slowly.	

## AN OSCILLATING PLATINUM WIRE

<b>TOPIC</b>	Catalysis 3	<b>DEMO #</b> 12.I.9
<b>REFERENCE</b>	Chemical Curiosities: Spectacular Experiments and Inspired Quotes, p. 260	
<b>EQUIPMENT</b>	platinum wire (about 15 cm) nichrome wire 500 mL flask bunsen burner and flint striker hot plate thermometer	
<b>CHEMICALS</b>	80 mL methanol	
<b>PROCEDURE</b>	<p>Form the platinum wire into a coil and attach it to a length of nichrome wire. Put 80 mL of methanol into the flask and use a hook arrangement so that the platinum is about 1 cm above the surface of the methanol. Remove the wire until the next step.</p> <p>Heat the methanol to between 40–60°C and quickly suspend the platinum wire over the methanol. Darken the room. The platinum will glow increasingly until the methanol catches fire and goes out after a few seconds. The platinum goes dark and then starts to glow again. The cycle repeats until the methanol has been used up. The reason the platinum increases its glow after the methanol flame has extinguished is because platinum is catalyzing the exothermic reaction between oxygen and methanol, and during the catalytic oxidation the heat produced is absorbed by the platinum.</p>	

## CATALYTIC OXIDATION OF ACETONE

<b>TOPIC</b>	Catalysis 4	<b>DEMO #</b> 12.I.10
<b>REFERENCE</b>	A Demo A Day: A Year of Chemical Demonstrations, p. 210	
<b>EQUIPMENT</b>	125 mL erlenmeyer flask heavy copper wire bunsen burner and flint striker penny pliers	
<b>CHEMICALS</b>	10 mL acetone	
<b>PROCEDURE</b>	<p>Encase the penny in wire and leave enough to form a hook over the lip of the flask. Pour 10 mL of acetone into the flask and arrange the wire so that the penny hangs just above, but not touching, the acetone surface.</p> <p>Using the pliers, heat the penny to red heat and quickly suspend the penny in the flask, just over the acetone. Turn out the lights. The penny continues to glow brightly until the acetone is exhausted. The penny has a golden colour because that is the true colour of copper: the reddish colour normally seen occurs because the surface of copper quickly oxidizes to <math>\text{Cu}_2\text{O}</math> in air. The great "ruby" in Queen Elizabeth's crown is actually cuprite, <math>\text{Cu}_2\text{O}</math>.</p>	

## AUTOCATALYSIS

<b>TOPIC</b>	Autocatalysis	<b>DEMO #</b> 12.I.11
<b>REFERENCE</b>	Chemical Demonstrations: A Sourcebook for Teachers, Volume 1, p. 105	
<b>EQUIPMENT</b>	2 – 100 mL beaker 100 mL graduated cylinder dropping pipet	
<b>CHEMICALS</b>	4.0 g of potassium chlorate 12.5 g of sodium sulphite 5 mg of bromophenol blue indicator 10 mL of 3 M sulphuric acid	
<b>PROCEDURE</b>	Dissolve 4.0 g of potassium chlorate, 12.5 g of sodium sulphite and 5 mg of bromophenol blue indicator in 50 mL of distilled water. To a second beaker add 4 mL of 3 M sulphuric acid in 50 mL of distilled water.	

Slowly, with stirring, add the acid in the second beaker to the contents of the first beaker. Stir until everything has dissolved, to produce a blue–violet colour. Fill a 100 mL graduated cylinder with the resulting solution.

Carefully add two droppers full of 3 M sulphuric acid to the top of the solution in the graduated cylinder. Shortly, a yellow colour appears at the top of the solution in the cylinder and a yellow–blue interface forms. Over a period of several minutes the yellow–blue interface moves down the cylinder.

The reaction occurring is  $\text{ClO}_3^- + 3\text{HSO}_3^- \rightarrow \text{Cl}^- + 3\text{SO}_4^{2-} + 3\text{H}^+$ .

This reaction requires the presence of acid as a catalyst and as the reaction continues more of the acidic catalyst is produced. The initial solution has a pH in the range 6.5–7.0 (bromophenol blue is blue). After addition of the sulphuric acid (turning the bromophenol yellow) the autocatalytic effect turns more and more of the solution acidic (yellow).

## INHIBITION OF HYDROGEN PEROXIDE

<b>TOPIC</b>	Reaction inhibitor	<b>DEMO #</b> 12.I.12
<b>REFERENCE</b>	A Demo A Day: A Year of Chemical Demonstrations, p. 192	
<b>EQUIPMENT</b>	1 L round bottom flask 2 dropping pipettes	
<b>CHEMICALS</b>	0.1 M iron (III) nitrate: $\text{Fe}(\text{NO}_3)_3 \cdot 9\text{H}_2\text{O}$ (4.04 g $\text{Fe}(\text{NO}_3)_3 \cdot 9\text{H}_2\text{O}$ / 100 mL) 30% hydrogen peroxide 0.1 M sodium phosphate tribasic: $\text{Na}_3\text{PO}_4 \cdot 10\text{H}_2\text{O}$ (3.44 g $\text{Na}_3\text{PO}_4 \cdot 10\text{H}_2\text{O}$ / 100 mL)	
<b>PROCEDURE</b>	Pour 25 mL of hydrogen peroxide into the flask. Add a few drops of iron (III) nitrate solution to the flask. A highly exothermic reaction quickly produces steam. After a few seconds: add several drops of sodium phosphate to immediately stop the reaction. (The inhibition effect is a result of precipitating the iron (III) ion with phosphate ion.)	

## CHEMILUMINESCENCE — THE FIREFLY REACTION

<b>TOPIC</b>	Exothermic reaction	<b>DEMO #</b> 12.I.13
<b>REFERENCE</b>	A Demo A Day: A Year of Chemical Demonstrations, p. 40	
<b>EQUIPMENT</b>	tall stand with clamps and clear Tygon™ tubing 2 L beaker big glass funnel small step ladder	
<b>CHEMICALS</b>	Solution A: Dilute 4.0 g anhydrous sodium carbonate ( $\text{Na}_2\text{CO}_3$ ) to 500 mL, add 0.2 g luminol and stir to dissolve. Add 24.0g sodium bicarbonate ( $\text{NaHCO}_3$ ), 0.5 g ammonium carbonate [ $(\text{NH}_4)_2\text{CO}_3 \cdot \text{H}_2\text{O}$ ] and 0.4 g copper (II) sulphate ( $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ ) and dissolve. Dilute to 1 L.  Solution B: 50 mL of 3% $\text{H}_2\text{O}_2$ diluted to 1 L.	
<b>PROCEDURE</b>	Pour equal amounts of bottles A and B into funnel, with lights off (adjust pouring rate to get maximum glow). Luminol is a special molecule that reacts to form a product having electrons in a highly excited energy level. As the molecule loses energy and its electrons fall down to a lower energy state, the energy lost is in the form of blue light – an exothermic reaction.	

## RACQUET BALL

<b>TOPIC</b>	Activation energy	<b>DEMO #</b> 12.I.14
<b>REFERENCE</b>	A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 72	
<b>EQUIPMENT</b>	racquet ball scissors sharp knife or X-acto knife	
<b>CHEMICALS</b>	—	
<b>PROCEDURE</b>	<p>Find the equatorial seam on the racquet ball and mark the seam with a pen. Slice through the seam and/or puncture the seam with the tip of a sharp knife (CARE!). Proceed to cut the ball in half. Use one half to make the model, and repeat with the second half later, if desired. Press on the outer end of the bowl to invert the bowl and create a hat-like appearance. Mark a starting point on the rim of the "hat" on the inner side and use the scissors to carefully trim off a thin section around the edge. The idea is to keep the hemisphere/hat as circular as possible while trimming off more and more off the rim. When you are approaching the correct amount to be removed (which might take 4 or 5 thin rings cut off) the hat shape will take very little push on the crown of the hat to cause it to suddenly snap back to the relaxed hemisphere shape. Each time a ring is cut off drop the ball in its hat form with the hat <b>RIM</b> facing <b>DOWN</b>, and the hat <b>CROWN</b> facing <b>UP</b>, from a height of about 40–100 cm above a hard flat surface (like a bench top or a floor). When sufficient rubber has been removed, the hat will spontaneously unflex when it hits the hard surface and will jump 100–150 cm into the air! If the ball spontaneously snaps into the bowl form while being trimmed (even if a full circle has not been completed), immediately stop and see if you have the desired effect when the hat form is dropped. If you remove too much the ball is ruined and you will need to start again with the second (reserve) half of the original ball. To fine-tune the ball, clip off small amounts until the hat form snaps into the hemisphere form when dropped from about 40 cm height.</p> <p>To use the ball, invert the hemisphere into the hat form and drop the hat from about 5 cm above the surface of a table or lab bench. Point out that the original potential energy possessed by the hat is converted into kinetic energy as it approaches the surface. At the moment of collision, the rubber is deformed and the kinetic energy is converted to and stored as potential energy by deforming and flexing the molecular structure. This potential energy is converted back into kinetic energy as the molecular structure flexes back to its original shape, as manifested by a small rebounding action. The remaining energy is lost to the surroundings.</p> <p>Drop the hat from about 10 cm and see essentially the same thing happen. Then drop the hat from about 40 cm and the hat suddenly jumps high into the air. This is analogous to the concept of activation energy: at some height the potential energy can be converted to sufficient kinetic energy to cause sufficient deformation of the rubber to complete the unflexing and the potential energy originally stored by flexing the hemisphere into the hat form is suddenly released in the form of a large amount of kinetic energy.</p>	

## ACTIVATED COMPLEX

<b>TOPIC</b>	Activated complex	<b>DEMO #</b> 12.I.15
<b>REFERENCE</b>	A Demo A Day: A Year of Chemical Demonstrations, p. 198	
<b>EQUIPMENT</b>	600 mL beaker 1 – 250 mL graduated cylinder 1 – 100 mL graduated cylinder 1 – 25 mL graduated cylinder hot plate watch glass to cover beaker thermometer	
<b>CHEMICALS</b>	150 mL 6% hydrogen peroxide: $\text{H}_2\text{O}_2$ 0.3 M cobalt (II) chloride: $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ (7.14 g $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ / 100 mL) 0.3 M sodium potassium tartrate: $\text{NaKC}_4\text{H}_4\text{O}_6 \cdot 4\text{H}_2\text{O}$ (42.3 g $\text{NaKC}_4\text{H}_4\text{O}_6 \cdot 4\text{H}_2\text{O}$ / 500 mL)	
<b>PROCEDURE</b>	Add 200 mL of sodium potassium tartrate and 65 mL of hydrogen peroxide to the beaker: cover with the watch glass and heat to $50^\circ\text{C}$ . When the temperature is correct: pour in 15 mL of the cobalt (II) chloride solution and replace the watch glass. The pink solution turns green (which is the colour of the activated complex) and when the reaction is finished the pink colour of the cobalt (II) ion catalyst reappears.	

## THE RATE OF FUNNELS

<b>TOPIC</b>	Rate analogy	<b>DEMO #</b> 12.I.16
<b>REFERENCE</b>	A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 230	
<b>EQUIPMENT</b>	2 – 2 L beakers 3 funnels: each with a different bore ring stand 3 rings (sized to hold the 3 funnels) 1000 mL graduated cylinder	
<b>CHEMICALS</b>	2 L of water with food colouring or indicator dye (colour not important)	
<b>PROCEDURE</b>	For each of the 3 funnels: pour water through them for 15 seconds and record the volume that passes. Calculate the # of mL/s for each.  Set up the funnels above each other so that they pour into the 1 L beaker. The medium funnel is at the top: the smallest in the middle and the largest at the bottom. Have students predict the “rate-determining” funnel and find the volume of water which flows in 15 seconds. Rearrange the funnels and repeat: the flow rate should be the same.	

## THORIUM OXIDE AS A CATALYST

<b>TOPIC</b>	Effect of catalyst on reaction mechanism	<b>DEMO #</b> 12.I.17
<b>REFERENCE</b>	Chemical Demonstrations: A Sourcebook for Teachers, Volume 1, p. 99	
<b>EQUIPMENT</b>	matches toothpick or straightened paper clip	
<b>CHEMICALS</b>	cigarette ashes 2 – sugar cubes	
<b>PROCEDURE</b>	<p>Impale a sugar cube on a toothpick or a straightened paper clip and hold the sugar cube over the flame of a match. The sugar cube will melt and caramelize slightly, but will not burn.</p> <p>Take another sugar cube and rub it generously with cigarette ashes on at least two sides. Hold the sugar cube in the flame of a match so that the sides covered by ashes are in the flame. The sugar will burn.</p> <p>Point out that tobacco plants are grown in areas containing very small amounts of naturally occurring uranium. A very small amount of this uranium in the soil is taken up into the tobacco plant. When uranium undergoes slow radioactive decay, one of its products is radioactive thorium and it is the oxide of this tiny amount of radioactive thorium contained in the ashes that acts as a catalyst for the burning of sugar. Point out how effective this catalyst is and note out that another decay product of uranium drawn into the lungs of smokers is radioactive radon gas.</p>	

## THE CHEMIST'S "RULE OF THUMB"

<b>TOPIC</b>	Doubling the reaction rate	<b>DEMO #</b> 12.I.18
<b>REFERENCE</b>	Chemical Demonstrations: A Sourcebook for Teachers, Volume 2, p. 141	
<b>EQUIPMENT</b>	hotplate thermometer 600 mL beaker 6-test tubes (18 x 150 mm)	
<b>CHEMICALS</b>	250 mL of 0.005 M $\text{KMnO}_4$ (dilute 0.20 g of $\text{KMnO}_4$ to 250 mL) 250 mL of saturated oxalic acid (about 15-17 g of oxalic acid/250 mL)	
<b>PROCEDURE</b>	<p>Put 25.0 mL of 0.005 M <math>\text{KMnO}_4</math> in each of 3 test tubes. Add 25.0 mL of saturated oxalic acid to each of 3 different test tubes. Half-fill a 600 mL beaker with tap water, put in all the test tubes and put the beaker on an operating hot plate (don't set the hot plate to too high a temperature, so as to keep the heating rate slow enough to allow the temperature of the test tube contents to closely track the temperature of the water in the beaker).</p> <p>When the temperature reaches 40°C, remove one tube of oxalic acid, note the time and quickly pour the oxalic acid into one tube of <math>\text{KMnO}_4</math> and pour back and forth two more times, finally placing the full tube back in the water bath. Stop timing when the test tube is yellow with no trace of pink colour. The difference between the initial and final times is the reaction time.</p> <p>When the temperature reaches 50°C, repeat the mixing procedure with the next set of test tubes and record the time. When the temperature reaches 60°C, repeat the mixing procedure with the last set of test tubes and record the time.</p>	

The following data can be expected:

Temperature	Time
40°C	100 s
50°C	49 s
60°C	24 s

## EQUILIBRIUM IS NOT FAIR

<b>TOPIC</b>	Equilibrium analogy	<b>DEMO #</b> 12.II.1
<b>REFERENCE</b>	A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 228	
<b>EQUIPMENT</b>	12 – glass tubes: 6 mm OD x 35 cm 12 – glass tubes: 10 mm OD x 35 cm 24 – 100 mL graduated cylinders	
<b>CHEMICALS</b>	2 L of water with food colouring or indicator dye (any colour)	
<b>PROCEDURE</b>	<b><i>Students to do this in pairs:</i></b> Take 100 mL of coloured water in one 100 mL cylinder — this is the REACTANT cylinder. Use the large diameter tube to transfer water from the REACTANT cylinder to the PRODUCT cylinder while simultaneously transferring water from the PRODUCT cylinder to the REACTANT cylinder (one transfer each way per “cycle”). [Transfer is accomplished by lowering the tube to the bottom of the cylinder and placing a finger over the top of the tube.] Continue until the volumes in the cylinders remain constant.	

## THE AUTOMATIC WATER FOUNTAIN: HELIUM EFFUSION

**TOPIC** Spontaneous movement of gases 1 **DEMO #** 12.II.2

**REFERENCE** Journal of Chemical Education, Vol. 61. No. 3, March 1984, p. 251

**EQUIPMENT** unglazed ceramic electrolysis cup  
one-holed rubber stopper to fit ceramic cup  
squeeze bottle for delivering liquids, with screw cap and separate side arm for liquid delivery  
one-hole rubber stopper to fit top of squeeze bottle  
8–10 cm length of glass tubing to fit above two rubber stoppers  
400 mL beaker

**CHEMICALS** tank of helium gas, with a length of rubber tubing attached

**PROCEDURE** Connect the two rubber stoppers at opposite ends of the glass tubing in such a way that one stopper is able to fit into the squeeze bottle and the unglazed ceramic cup can be connected in an inverted position to the rubber stopper at the other end of the tubing.

Fill the squeeze bottle with water to the fill line marked on the bottle, insert the rubber stopper and connect the ceramic cup above.

Invert the 400 mL beaker and fill the beaker from below with helium. Place the inverted beaker over the ceramic cup. Water should pour out of the delivery tube on the bottle. Remove the beaker, keeping it inverted, and the water flow should stop. Replace the beaker over the ceramic cup and water should again flow out. Remove the beaker, hold it upright and wave it around a few times to fill it with air. Invert the beaker again, hold it over the ceramic cup and no water flows out.

**Explanation:** When the cup is first connected to the squeeze bottle, molecules in the air are able to pass inward and outward through the pores in the ceramic at an equal rate so that no pressure change occurs. When the outside atmosphere is helium, the air molecules inside the cup effuse outward through the pores in the ceramic slower than the helium effuses into the cup. As a result, the pressure inside the cup increases and water is pushed out of the connecting bottle. When the beaker filled with helium is removed, the balance between the gases is restored and the pressures inside and out become equal again.

**Variation:** Fill the 400 mL beaker with carbon dioxide gas and bubbles of air will be sucked into the squeeze bottle because air leaves the ceramic cup faster than the heavy and slow-moving carbon dioxide molecules can enter.

**Interesting fact:** Creating a “super-hard” vacuum on earth is virtually impossible because the tiny amount of helium naturally present in the atmosphere can diffuse through several inches of high-density steel faster than even the best vacuum pumps are able to remove the helium.

## SHRINKING SUDS

**TOPIC** Spontaneous movement of gases 2 **DEMO #** 12.II.3

**REFERENCE** Twenty Demonstrations Guaranteed to Knock Your Socks Off! , p. 50

**EQUIPMENT** 2 L graduated cylinder  
dewar flask (to contain dry ice)  
dropping pipet

**CHEMICALS** several chunks of dry ice  
dish detergent (Joy® or Dawn® work best)

**PROCEDURE** Fill the graduated cylinder one third full of warm water. Drop in some dry ice and observe for a while. Add a squirt (2–3 mL) of detergent. A large number of bubbles form and overflow the cylinder. Scoop up a handful of the bubbles and watch them rapidly shrink from golf ball size to BB size.

The foggy CO<sub>2</sub> bubbles shrink quickly because CO<sub>2</sub> is relatively soluble in water and can rapidly diffuse through the soapy membrane of the bubbles. The bubbles do not shrink in the cylinder because the atmosphere inside the cylinder is CO<sub>2</sub>, which diffuses into the bubbles as fast as it diffuses out. The fogginess inside the bubbles is due to condensing water vapour.

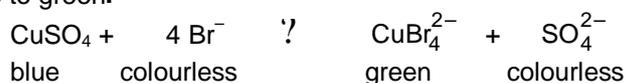
## ENDOTHERMIC REACTION BETWEEN $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$ AND $\text{NH}_4\text{SCN}$

<b>TOPIC</b>	Entropy-driven reaction	<b>DEMO #</b> 12.II.4
<b>REFERENCE</b>	Chem 13 News, November 1976, p. 9	
<b>EQUIPMENT</b>	small, flat unvarnished/unpainted wooden block (about 3" x 4" x 1") 250 mL beaker stirring rod	
<b>CHEMICALS</b>	about 32 g of barium hydroxide, $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$ about 16 g of ammonium thiocyanate, $\text{NH}_4\text{SCN}$ water	
<b>PROCEDURE</b>	<p>Place a few millilitres of water on top of the block and place the beaker on top of the wet wood. (The idea is to have a layer of liquid between the beaker bottom and the wood. If this is done out of sight of the class, the final surprise effect is enhanced.)</p> <p>Dump about 32 g of barium hydroxide into the beaker, followed by about 16 g of ammonium thiocyanate. Stir the mixture for about one minute, during which time the mixture liquifies* and cools to about <math>-55^\circ\text{C}</math>. If the beaker is now lifted, the block of wood is frozen to it.</p> <p>* NOTE: Students understand that ice requires the addition of heat in order to melt but they frequently have the idea that "real" chemical reactions only give off heat. The fact that the mixture liquifies tends to make them think the mixture has given off so much heat that the solids have MELTED. Lead them "down the garden path" for a moment by suggesting it is reasonable to assume chemicals may absorb energy if heat is PROVIDED, but chemicals will NOT somehow go up a huge energy hill on their own because the temperature would have to drop dramatically. Picking up the block provides a "discrepant event" that forces students to re-think their assumptions.</p>	

## EQUILIBRIUM ON AN OVERHEAD

- TOPIC** Le Chatelier's Principle 1 **DEMO #** 12.II.5
- REFERENCE** A Demo A Day: A Year of Chemical Demonstrations, p. 220. Modified by Jim Hebden
- EQUIPMENT** light box  
3 – 400 mL beakers  
3 – stirring rods  
100 mL graduated cylinder
- CHEMICALS** 500 mL of 0.25 M copper (II) sulphate (31.2 g  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  / 500 mL)  
30 g of ammonium bromide,  $\text{NH}_4\text{Br}$ , divided into two 15 g portions  
20 g of ammonium sulphate,  $(\text{NH}_4)_2\text{SO}_4$
- PROCEDURE** Put 100 mL of copper (II) sulphate solution in each of three 400 mL beakers. Place the beakers on the light box and use one beaker as a reference.

To each of the other two beakers add 15 g of  $\text{NH}_4\text{Br}$  and stir to dissolve. The colour shifts from blue to green.



To one of the green solutions add 50 mL of ammonium sulphate and stir (not all the solid may dissolve). The colour shifts back to blue.

## THE EFFECT OF PRESSURE CHANGES ON AN EQUILIBRIUM I

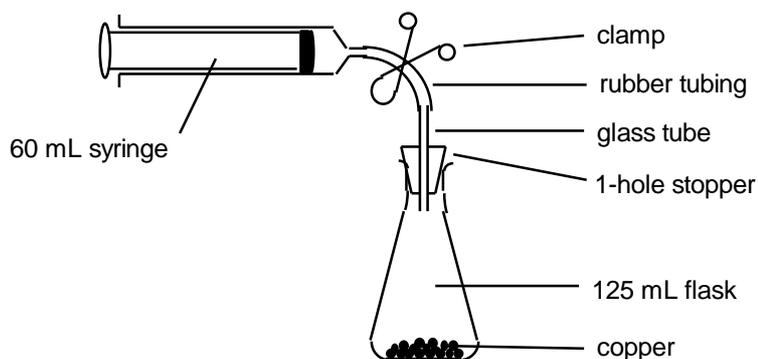
TOPIC Le Chatelier's Principle 2 DEMO # 12.II.6

REFERENCE commonly known

EQUIPMENT 125 mL erlenmeyer flask, with one-hole rubber stopper to fit flask  
60 mL plastic syringe, without tip  
10 cm length of 8 mm OD glass tubing  
10 cm length of latex tubing (to fit over glass tubing at one end and syringe tip at other end)  
tubing clamp

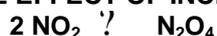
CHEMICALS 1 g of copper (wire, sheet, filings or "shot")  
5 mL of concentrated nitric acid

PROCEDURE Fit a 125 mL flask with a one-hole stopper, through which is put a 10 cm length of 8 mm OD pyrex tubing (use a few drops of glycerine to aid in pushing the tubing into the rubber stopper; protect your hands to prevent injury in case the glass tube breaks). Place about 1 g of copper metal (wire, sheet, filings or "shot") into the flask. Place a clamp over the middle of a 10 cm piece of latex tube. Push one end of the latex tube onto the end of a 60 mL syringe and push the other end of the latex tube over the top of the glass tube protruding from the flask. The final setup should resemble the diagram below.



**DO THE FOLLOWING IN AN OPERATING FUME HOOD.** Obtain about 5 mL of concentrated nitric acid. Remove the rubber tubing from the top of the glass tube. Remove the rubber stopper from the flask, pour in the nitric acid and re-stopper the flask. The reaction between nitric acid and copper creates substantial amounts of red-brown  $\text{NO}_2(\text{g})$ . Swirl the flask to help the gas production. When the flask is filled with  $\text{NO}_2$ , place the rubber tube back onto the glass tube in the flask and open the clamp. Slowly draw back the plunger in the syringe to fill the syringe with red-brown gas. When full, re-clamp the latex tubing and remove the tubing from the glass tube.

### DEMONSTRATING THE EFFECT OF INCREASED PRESSURE ON THE EQUILIBRIUM



Hold the rubber tube firmly onto the end of the syringe while quickly and firmly pushing in the plunger. The colour gets darker and then become lighter again after a few seconds. Pull out the plunger again to reverse the effect. The effect is not highly dramatic, but is very noticeable and demonstrates the desired effect quite nicely.

**What is Happening:** When the pressure increases, the concentration of both  $\text{NO}_2(\text{g})$  and  $\text{N}_2\text{O}_4(\text{g})$  increases inside the syringe.  $\text{N}_2\text{O}_4$  is colourless, so the effect of increasing its concentration can't be seen, but  $\text{NO}_2$  is red-brown and **its increased concentration as a result of the increased pressure is seen as a darker red-brown colour**. This increased pressure causes the equilibrium to shift and use up gas molecules, lowering the pressure. Since 2 molecules of  $\text{NO}_2$  react to produce 1 molecule of  $\text{N}_2\text{O}_4$ , the increased forward reaction uses up red-brown  $\text{NO}_2$  molecules and replaces them with colourless  $\text{N}_2\text{O}_4$  molecules, giving rise to the observed decrease in colour intensity.

## THE EFFECT OF PRESSURE CHANGES ON AN EQUILIBRIUM II

**TOPIC** Le Chatelier's Principle 3 **DEMO #** 12.II.7

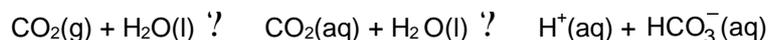
**REFERENCE** Chemical Activities: Teacher Edition, p. 8

**EQUIPMENT** 60 mL plastic syringe  
small piece of putty or plasticine  
400 mL beaker

**CHEMICALS** 200 mL of club soda or water saturated with carbon dioxide  
1 mL of universal indicator

**PROCEDURE** Put about 200 mL of club soda into a 400 mL beaker. Add several drops of universal indicator, to produce a definite coloration.

Fill the syringe half full of club soda by placing the open end down into the soda and pulling up on the plunger. Invert the plunger so as to have the small open end up and depress the plunger to expel the air. Place a piece of plasticine over the small end, to seal the tip, and pull out the plunger. Bubbles of carbon dioxide gas will form throughout the solution as the pressure is reduced.



By Le Chatelier's principle, the reduction of pressure causes the dissolved  $\text{CO}_2(\text{aq})$  to go into the gas phase. The loss of carbon dioxide from solution causes a loss of hydrogen carbonate ion and a decrease in the concentration of hydrogen ion, increasing the pH.

## GASES HAVE IT BACKWARDS

**TOPIC** Effect of temperature on gas solubility **DEMO #** 12.II.8

**REFERENCE** A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 186

**EQUIPMENT** 3 – 250 mL erlenmeyer flasks  
1 – rubber stopper (solid) to fit flask  
2 – rubber stoppers (2-hole) to fit flasks  
2 – 50 cm glass tubes (8 mm OD)  
2 – 15 cm glass tubes (8 mm OD)  
2 – latex tubing to fit glass tubes (5 cm length)  
2 – pinch clamps  
2 – syringes, with needles (20 mL)  
2 – 1 L beakers  
bunsen burner  
stand with ring and wire mesh pad

**CHEMICALS** tank of carbon dioxide  
ice

**PROCEDURE** Fit each stopper with glass tubes such that the 50 cm tube extends almost to the bottom of the flask and the 15 cm tube is about 1 cm from the bottom. Fit a latex tube and pinch clamp to the top of the 15 cm tube.

Fill one flask completely with ice water and place in a 1 L beaker containing ice water.  
Fill the other flask with  $>80^{\circ}\text{C}$  water and place in 1 L beaker containing  $>80^{\circ}\text{C}$  water.

Fill an empty flask with carbon dioxide and fill the two syringes with gas. Inject 20 mL of gas through the latex tubing of each flask. More gas is absorbed by the ice water and less is absorbed by the hot water, as shown by the volumes of water displaced into the long glass tubes. Therefore, the gas is more soluble in cold water.

## OSTWALD OXIDATION OF AMMONIA

**TOPIC** The Haber and Ostwald reactions **DEMO #** 12.II.9

**REFERENCE** A Demo A Day: A Year of Chemical Demonstrations, p. 224

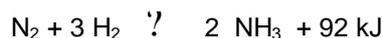
**EQUIPMENT** 500 mL filter flask  
bunsen burner and flint striker  
15 cm platinum wire  
nichrome wire

**CHEMICALS** concentrated ammonia

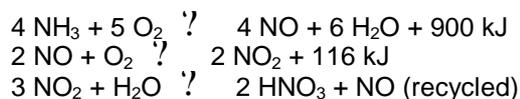
**PROCEDURE** Put 100 mL of ammonia in the flask. Coil the platinum wire and suspend it with the nichrome wire just above the surface of the ammonia: form a hook at the top of the nichrome wire to maintain the position of the platinum above the ammonia. Remove the wire, heat the platinum in a burner flame and quickly replace the wire in the flask. Darken the room. The platinum glows red hot. If bursts of oxygen are introduced into the flask, jets of flame are produced in a cyclic fashion.

The Haber process is used to make ammonia and the Ostwald process is then used to convert the ammonia to nitrates.

The reaction for the Haber process is:



The overall series of reactions for the Ostwald process is:



## A GOLDEN RAIN OF LEAD (II) IODIDE CRYSTALS

<b>TOPIC</b>	Crystallization reactions	<b>DEMO #</b> 12.III.1
<b>REFERENCE</b>	Chemical Curiosities: Spectacular Experiments and Inspired Quotes, p. 5	
<b>EQUIPMENT</b>	light box black cardboard backdrop 1 L florence flask, with stopper	
<b>CHEMICALS</b>	Lead (II) acetate, 5.6 g sodium iodide, 5.2 g concentrated acetic acid, 5 mL 800 mL of hot (70–80°C) water	
<b>PROCEDURE</b>	<p>Dissolve about 5.6 g of lead (II) acetate in 100 mL water and separately about 5.2 g of sodium iodide in 100 mL of water. Mix the solutions, let settle and filter off.</p> <p>Add about 5 mL of acetic acid to 800 mL of hot water in the flask and then about 3 g of lead (II) iodide, to make a homogeneous, saturated and almost colourless solution. Swirl the flask slowly and cool it under cold water. A few seconds later, beautiful golden yellow crystals form throughout the solution and fill the flask. The crystallization process is best seen on a light box against a black background with room lights off. Spectacular!</p>	

## FRACTIONAL CRYSTALLIZATION

<b>TOPIC</b>	Solubility diagrams	<b>DEMO #</b> 12.III.2
<b>REFERENCE</b>	A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 136	
<b>EQUIPMENT</b>	250 mL beaker 400 mL beaker hot plate or bunsen burner and flint striker hot hand or beaker tongs 25 x 150 mm test tube test tube clamp dropping pipet ice	
<b>CHEMICALS</b>	sodium chloride, 20 g potassium nitrate, 20 g acetone in wash bottle 5 mL of 0.1 M lead (II) nitrate, $\text{Pb}(\text{NO}_3)_2$	
<b>PROCEDURE</b>	<p>Add 20 g of sodium chloride, 20 g of potassium nitrate and 80 mL of distilled water to a 250 mL beaker. Heat the solution, with stirring, until the volume is reduced by one half. As boiling occurs, crystals of both sodium chloride and potassium nitrate are formed.</p> <p>Remove the beaker from the heat source and decant the solution into the test tube. Cool the test tube in an ice bath. As the solution cools, crystals of potassium nitrate form.</p> <p>When the test tube has cooled, decant the solution from the test tube, leaving the crystals behind. Wash the crystals with about 5 mL of acetone and discard the acetone. Dissolve the crystals in a little distilled water. To test for the presence of chloride ions, add a few drops of lead nitrate solution. If chloride is present, a white precipitate forms.</p> <p>Note: The solubility of sodium chloride is about 36 g/100 mL at 10°C and 40 g/100 mL at 100°C. The solubility of potassium chloride is 20 g/100 mL at 10°C and 230 g/100 mL at 100°C.</p>	

## SOLUBILITY OF CALCIUM ACETATE

<b>TOPIC</b>	Solubility exception 1	<b>DEMO #</b> 12.III.3
<b>REFERENCE</b>	A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 188	
<b>EQUIPMENT</b>	125 mL erlenmeyer flask hot plate thermometer 600 mL beaker (for ice bath) ice stirring rod Hot Hand™ or insulated glove	
<b>CHEMICALS</b>	100 mL of saturated calcium acetate (dissolve 37.4 g of calcium acetate in 100 mL of cold distilled water)	
<b>PROCEDURE</b>	Place 100 mL of saturated calcium acetate in a 125 mL erlenmeyer flask. The calcium acetate solution should be clear. Place a thermometer in the flask of calcium acetate solution and heat it on a hot plate to about 85°C. The solution should become cloudy as calcium acetate crystallizes from solution. Remove the flask and place it in an ice bath (keep adding ice if necessary). Stir the solution as it cools. By the time the solution has chilled to 15°C the crystals should have redissolved.	

## THE EFFECT OF TEMPERATURE ON SOLUBILITY

**TOPIC** Solubility exception 2 **DEMO #** 12.III.4

**REFERENCE** Chemical Demonstrations: A Sourcebook for Teachers, Volume 1, p. 35

**EQUIPMENT** 25 x 200 mm test tube  
stopper to fit test tube  
2 – 1 L beakers  
ice  
thermometer  
hot plate

**CHEMICALS** 130 g manganese (II) sulphate heptahydrate,  $\text{MnSO}_4 \cdot 7\text{H}_2\text{O}$  or 70 g manganese (II) sulphate,  $\text{MnSO}_4$   
a few drops of concentrated sulphuric acid

**PROCEDURE** Heat 100 mL of distilled water to  $27^\circ\text{C}$ , add a few drops of concentrated sulphuric acid and either 130 g of manganese sulphate heptahydrate,  $\text{MnSO}_4 \cdot 7\text{H}_2\text{O}$  or 70 g of manganese sulphate,  $\text{MnSO}_4$ . Decant the clear solution into a large test tube and seal with a rubber stopper.

Prepare an ice bath in a 1 L beaker and cool the mixture in the test tube. Pink crystals should form. Remove the test tube and heat it in a warm water bath on a hot plate until the temperature reaches  $27^\circ\text{C}$ . At this point the solution should have become clear again. (If the solution is not clear, shake the tube to help the dissolving process.) Continue heating the test tube and observe the formation of a white precipitate.

The pink crystals are  $\text{MnSO}_4 \cdot 7\text{H}_2\text{O}$  and the white precipitate is  $\text{MnSO}_4 \cdot \text{H}_2\text{O}$ .

## NAME THAT PRECIPITATE

<b>TOPIC</b>	Precipitation equations	<b>DEMO #</b> 12.III.5
<b>REFERENCE</b>	Chemical Demonstrations: A Sourcebook for Teachers, Volume 2, p. 123	
<b>EQUIPMENT</b>	10 – 18 x 150 mm test tubes test tube rack light box	
<b>CHEMICALS</b>	mercury (II) chloride, $\text{HgCl}_2$ sodium carbonate, $\text{Na}_2\text{CO}_3$ lead (II) nitrate, $\text{Pb}(\text{NO}_3)_2$ potassium iodide, $\text{KI}$ barium chloride, $\text{BaCl}_2$ sodium sulphate, $\text{Na}_2\text{SO}_4$ copper (II) chloride, $\text{CuCl}_2$	
<b>PROCEDURE</b>	Create separate dilute solutions by dissolving about 1 g of each chemical in 50 mL of distilled water. Fill each labelled test tubes about one-third full as follows.	

Tube 1 =  $\text{HgCl}_2$   
Tube 2 =  $\text{Na}_2\text{CO}_3$   
Tube 3 =  $\text{Pb}(\text{NO}_3)_2$   
Tube 4 =  $\text{KI}$   
Tube 5 =  $\text{HgCl}_2$   
Tube 6 =  $\text{KI}$   
Tube 7 =  $\text{BaCl}_2$   
Tube 8 =  $\text{Na}_2\text{SO}_4$   
Tube 9 =  $\text{CuCl}_2$   
Tube 10 =  $\text{Na}_2\text{CO}_3$

Pour tube 1 into tube 2. Ask students to write down the formula equation for the reaction and suggest the identity of the precipitate. Once the identity of the precipitate is established, students should write the complete and net ionic equations for the reaction. Repeat by pouring tube 3 into 4, tube 5 into 6, tube 7 into 8, and tube 9 into 10.

## SOLUBILITY PRODUCT CONSTANT, $K_{SP}$

<b>TOPIC</b>	Will a precipitate form?	<b>DEMO #</b> 12.III.6
<b>REFERENCE</b>	A Demo A Day: A Year of Chemical Demonstrations, p. 146	
<b>EQUIPMENT</b>	2 – 50 mL burettes light box	
<b>CHEMICALS</b>	100 mL of 0.01 M $Pb(NO_3)_2$ (0.331 g of lead (II) nitrate diluted to 100 mL) 100 mL of 0.1 M KI (1.66 g of potassium iodide diluted to 100 mL)	
<b>PROCEDURE</b>	On a light box place a 50 mL burette filled with 0.01 M lead (II) nitrate. Put some KI solution into a second burette and position the second burette over the first burette. To the lead (II) nitrate solution add one drop of 0.1 M potassium iodide. When the precipitate disappears, add another drop. The point of this demonstration is to estimate as closely as possible how far down the burette the precipitate of $PbI_2$ goes until it dissolves.	

Assuming one drop is 0.05 mL and that the precipitate dissolves by the time it reaches the 6 mL mark in the lead (II) nitrate burette, the following calculation can be undertaken.

$$[I^-] = 0.1 \text{ M} \times \frac{0.05 \text{ mL}}{6 \text{ mL}} = 8.3 \times 10^{-4} \text{ M}$$

$$[Pb^{2+}] = 0.01 \text{ M} \times \frac{1 \text{ mL}}{1 \text{ mL}} = 0.01 \text{ M}$$

$$K_{sp} = [Pb^{2+}][I^-]^2 = 7 \times 10^{-9}$$

The accepted value of  $K_{sp}$  for  $PbI_2$  is  $8.5 \times 10^{-9}$ .

## CONCENTRATION DEPENDENCE OF A REACTION

**TOPIC** Water hardness **DEMO #** 12.III.7

**REFERENCE** A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 192

**EQUIPMENT** Light box  
2 – 250 mL beakers  
2 – 100 mL graduated cylinders  
2 – 10 mL graduated cylinders

**CHEMICALS** 101 mL of limewater (saturated calcium hydroxide,  $\text{Ca}(\text{OH})_2$ )  
101 mL of club soda (saturated carbonic acid solution) or carbon dioxide bubbled through water to produce a saturated solution

**PROCEDURE** To the first 250 mL beaker add 100 mL of lime water and 1 mL of club soda. A precipitate forms.

To the second 250 mL beaker add 100 mL of club soda and 1 mL of lime water. No precipitate forms.

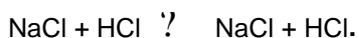
In the first beaker, the  $[\text{Ca}^{2+}]$  is about  $2.1 \times 10^{-3}$  M and the  $[\text{OH}^-]$  is about  $4.3 \times 10^{-3}$  M. In this basic solution, the added carbonic acid is primarily in the form of carbonate ion. Since saturated carbon dioxide at  $25^\circ\text{C}$  has a solubility of 1.45 g/L, the carbonate ion is about 0.033 M. When diluted by the added limewater, the carbonate ion has a molarity of about  $3.3 \times 10^{-4}$  M. The trial ion product for calcium and carbonate ions is therefore about  $7 \times 10^{-7}$ . Since  $K_{\text{sp}}$  for calcium carbonate is  $5 \times 10^{-9}$ , a precipitate forms.

In the second beaker, the  $[\text{Ca}^{2+}]$  from the lime water is about  $2.1 \times 10^{-5}$  M. Since the molarity of the carbon dioxide is about 0.033 M, and very little  $\text{CO}_2$  dissociates to form hydrogen carbonate ion, and very little of the hydrogen carbonate ion dissociates to carbonate ion, there is insufficient carbonate ion for a precipitate to form with the calcium ions present.

## COMMON ION EFFECT

<b>TOPIC</b>	Common ion effect 1	<b>DEMO #</b> 12.III.8
<b>REFERENCE</b>	A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 232 (modified)	
<b>EQUIPMENT</b>	250 mL beaker	
<b>CHEMICALS</b>	about 100 mL of saturated sodium chloride solution about 50 mL of concentrated hydrochloric acid	
<b>PROCEDURE</b>	Pour about 100 mL of saturated sodium chloride into a 250 mL beaker and then add about 50 mL of concentrated hydrochloric acid. A white precipitate forms immediately.	

Ask students what the precipitate is. They will usually assume the reaction is a double replacement and then get stuck:



Eventually they will realize that the only possible formula for the precipitate is NaCl. Add about 50 mL of water to the mixture and the precipitate dissolves.

Saturated sodium chloride is about 6 M and the addition of 12 M chloride ion causes the equilibrium



to shift to the reactants side.

## COMMON ION EFFECT REVISITED

<b>TOPIC</b>	Common ion effect 2	<b>DEMO #</b> 12.III.9
<b>REFERENCE</b>	A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 233	
<b>EQUIPMENT</b>	250 mL beaker	
<b>CHEMICALS</b>	100 mL of saturated potassium chlorate 5 mL of saturated potassium chloride	
<b>PROCEDURE</b>	Place 100 mL of saturated potassium chlorate solution into a 250 mL beaker. Upon adding 5 mL of saturated potassium chloride to the beaker, shining rhombic crystals of potassium chlorate are formed.	

## THE AMPHIPROTIC BEHAVIOUR OF $\text{Al}(\text{OH})_3$

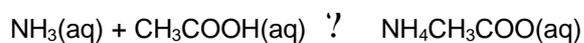
<b>TOPIC</b>	Amphiprotic ions	<b>DEMO #</b> 12.IV.1
<b>REFERENCE</b>	Chemical Curiosities: Spectacular Experiments and Inspired Quotes, p. 137	
<b>EQUIPMENT</b>	1 L beaker magnetic stirrer, with stir bar white card stock, to serve as a background dropper	
<b>CHEMICALS</b>	250 mL of 6 M NaOH 30 g of finely powdered aluminum potassium sulphate (potassium alum) 50 mL of concentrated sulphuric acid universal indicator 750 mL of distilled water	
<b>PROCEDURE</b>	<p>Dissolve 30 g of potassium alum in 750 mL of distilled water and stir with a magnetic stirrer. Add 5 mL of concentrated sulphuric acid to the mixture and then a few drops of universal indicator to give a red colour to the solution. Using a dropper, add small amounts of 6 M NaOH to the stirred mixture until the solution becomes milky as <math>\text{Al}(\text{OH})_3</math> precipitates. Further addition of NaOH causes the milkiness to suddenly disappear and become clear again. If desired, sulphuric acid can be added dropwise to reverse the reaction sequence.</p> <p>The addition of aluminum ion to an acidic solution produces <math>[\text{Al}(\text{H}_2\text{O})_6]^{3+}</math>. Addition of <math>\text{OH}^-</math> produces the following sequence of ions: <math>[\text{Al}(\text{H}_2\text{O})_5(\text{OH})]^{2+}(\text{aq})</math>, <math>[\text{Al}(\text{H}_2\text{O})_4(\text{OH})_2]^+(\text{aq})</math>, <math>[\text{Al}(\text{H}_2\text{O})_3(\text{OH})_3](\text{s})</math>, <math>[\text{Al}(\text{H}_2\text{O})_2(\text{OH})_4]^-</math> (aq), <math>[\text{Al}(\text{H}_2\text{O})(\text{OH})_5]^{2-}(\text{aq})</math> and <math>[\text{Al}(\text{OH})_6]^{3-}(\text{aq})</math>. The <math>[\text{Al}(\text{H}_2\text{O})_3(\text{OH})_3](\text{s})</math> is commonly labelled "<math>\text{Al}(\text{OH})_3</math>".</p>	

## RELATIVE ACID STRENGTH

<b>TOPIC</b>	Strong vs weak acids	# 12.IV.2
<b>REFERENCE</b>	A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 190	
<b>EQUIPMENT</b>	pH meter or pH paper 3 – 250 mL beakers conductivity meter	
<b>CHEMICALS</b>	3 – 5 cm strips of magnesium ribbon 100 mL of 1.0 M hydrochloric acid 100 mL of 1.0 M oxalic acid 100 mL of 1.0 M acetic acid	
<b>PROCEDURE</b>	Put 100 mL of each of 1.0 M acid in a separate labelled beaker. Inform students that all the acids are 1.0 M.  For each acid solution measure (a) the pH, (b) the conductivity, and (c) the rate at which a piece of magnesium ribbon reacts with the acid.  Hydrochloric acid has the lowest pH, the greatest conductivity and the greatest rate of reaction with magnesium. Acetic acid has the highest pH, the least conductivity and the slowest rate of reaction with magnesium.	

## CONDUCTIVITY OF A MIXTURE OF NH<sub>3</sub> AND CH<sub>3</sub>COOH

<b>TOPIC</b>	Weak acids and bases 1	<b>DEMO #</b> 12.IV.3
<b>REFERENCE</b>	A Demo A Day: A Year of Chemical Demonstrations, p. 147	
<b>EQUIPMENT</b>	conductivity apparatus 3 – 250 mL beakers stirring rod	
<b>CHEMICALS</b>	250 mL of 1 M acetic acid 250 mL of 1 M ammonium hydroxide distilled water	
<b>PROCEDURE</b>	Show students that the acetic acid solution contains very few ions. Rinse the electrodes with distilled water and show that the ammonium hydroxide solution contains very few ions. Pour 125 mL of the acetic acid into a clean, dry beaker and add 125 mL of ammonium hydroxide to the same beaker. Stir and show that the resulting mixture of ammonium acetate contains a substantial number of ions.	



## ETCHING GLASS WITH HF

<b>TOPIC</b>	Weak acids and bases 2	<b>DEMO #</b> 12.IV.4
<b>REFERENCE</b>	A Demo A Day: A Year of Chemical Demonstrations, p. 190	
<b>EQUIPMENT</b>	piece of glass or microscope slide (or beaker to be personalized) paraffin wax fume hood plastic petri plate, with lid chemical resistant gloves hot plate tongs large needle, nail or other scraper to incise wax	
<b>CHEMICALS</b>	concentrated hydrofluoric acid or 10 g of calcium fluoride and 15 mL of concentrated sulphuric acid	
<b>PROCEDURE</b>	<p>Melt the wax and coat the piece of glass with a thin layer of wax. After the wax has cooled, scrape away wax from the glass in such a way as to create a design or message.</p> <p>Place 10 g of calcium fluoride into a plastic petri plate. Slowly add 15 mL of concentrated sulphuric acid to the dish. Use tongs to place the glass to be etched into the mixture and cover with a plastic lid. After 15–20 minutes, remove the glass and rinse thoroughly with tap water. When the wax is removed with hot water, the message etched in the glass is revealed.</p> <p><b>CAUTION:</b> Hydrofluoric acid causes severe burns.</p>	

## PREPARATION AND PROPERTIES OF CARBON DIOXIDE

<b>TOPIC</b>	Acidity of CO <sub>2</sub> (aq)	<b>DEMO #</b> 12.IV.5
<b>REFERENCE</b>	A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 101	
<b>EQUIPMENT</b>	250 mL flask rubber tubing to deliver carbon dioxide into flask fritted glass gas bubbler (optional). Fish tank bubblers work well.	
<b>CHEMICALS</b>	source of carbon dioxide universal indicator	
<b>PROCEDURE</b>	Bubble carbon dioxide through about 100 mL of distilled water, until the solution is saturated. Test the pH of the solution by adding some universal indicator. The solution is acidic.	

## EFFECT OF TEMPERATURE ON THE $\text{NH}_3 / \text{NH}_4^+$ EQUILIBRIUM

**TOPIC** Effect of temperature on  $K_b$  **DEMO # 12.IV.6**

**REFERENCE** Chemical Demonstrations: A Sourcebook for Teachers, Volume 1, p. 87

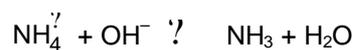
**EQUIPMENT** 600 mL beaker  
25 x 200 mm test tube  
cold water tap and sink  
bunsen burner and flint striker

**CHEMICALS** phenolphthalein solution  
a few millilitres of concentrated ammonium hydroxide

**PROCEDURE** To a 600 mL beaker add 400 mL of water, a few drops of phenolphthalein solution and ONE drop of concentrated ammonium hydroxide. The solution should be a light pink. Fill a 25 x 200 mm test tube half full with the coloured solution.

Heat the test tube over a bunsen burner. The colour should fade and disappear. Cooling the test tube under a stream of cold water restores the colour.

Ask student whether the equilibrium



is endothermic or exothermic.

## pH RAINBOW TUBE

<b>TOPIC</b>	The pH scale	<b>DEMO #</b> 12.IV.7
<b>REFERENCE</b>	Twenty Demonstrations Guaranteed to Knock Your Socks Off! , p. 4	
<b>EQUIPMENT</b>	clear glass or plastic tube, 1.5–3.0 cm ID and 40–80 cm long 2 – solid rubber stoppers to fit tube beaker having a volume exceeding that of the tube 2 – plastic pipettes	
<b>CHEMICALS</b>	distilled water universal indicator about 10 mL of 0.1 M hydrochloric acid about 10 mL of 0.1 M sodium hydroxide	
<b>PROCEDURE</b>	Stopper one end of the tube. Fill the tube with a 5% by volume solution of universal indicator in distilled water. Add a few drops of 0.1 M HCl to the tube, hold the stopper in the opposite end firmly and securely stopper the tube. Invert the tube, remove the top stopper (formerly the bottom), add an equal number of drops of 0.1 M NaOH and again securely stopper. Invert the tube and observe the full indicator spectrum spread throughout the tube.	

## ACIDIC AND BASIC SALTS

<b>TOPIC</b>	Hydrolysis 1	<b>DEMO #</b> 12.IV.8
<b>REFERENCE</b>	Chemical Curiosities: Spectacular Experiments and Inspired Quotes, p.135	
<b>EQUIPMENT</b>	2 – 18 x 150 mm test tubes, with stoppers 2 — 125 mL erlenmeyer flasks, with stoppers bunsen burner with flint striker clamp to hold flask in burner flame protective gloves 1 L beaker containing crushed ice, for an ice bath (optional)	
<b>CHEMICALS</b>	alcoholic phenolphthalein solution (0.050 g in 100 mL ethanol — do not use WATER) alcoholic methyl red solution (0.020 g in 100 mL ethanol — do not use WATER) 18 g of sodium acetate trihydrate, $\text{CH}_3\text{COONa}\cdot 3\text{H}_2\text{O}$ 15 g of magnesium chloride hexahydrate, $\text{MgCl}_2\cdot 6\text{H}_2\text{O}$ distilled water	
<b>PROCEDURE</b>	<p>Add 3 g of sodium acetate to a test tube and add 10 mL of distilled water. Add 5 drops of alcoholic phenolphthalein solution, stopper the test tube and shake and show that the solution is pale pink: a basic solution.</p> <p>Put 15 g of sodium acetate in a 125 mL flask and add 5 drops of alcoholic phenolphthalein solution. Stopper the colourless mixture, clamp the flask and heat it carefully over a burner flame. The sodium acetate melts quickly and produces a raspberry red solution. Cool the mixture under cold running water (or in an ice bath); when the mixture has solidified, the colour disappears.</p> <p>Add 3 g of magnesium chloride to a test tube and add 10 mL of distilled water. Add 5 drops of alcoholic methyl red solution, stopper the test tube, shake and show that the solution is bright red: an acidic solution.</p> <p>Put 12 g of magnesium chloride in a 125 mL flask and add 5 drops of alcoholic methyl red solution. Stopper the mixture, clamp the flask and heat it carefully over a burner flame. The magnesium chloride melts quickly and produces a bright red solution. Cool the mixture under cold running water (or in an ice bath); when the mixture has solidified, the colour disappears.</p> <p>Melting the sodium acetate allows water molecules to take part in the equilibrium <math display="block">\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-</math> The mobile ions react with the phenolphthalein. When the sodium acetate solidifies, the crystal lattice excludes the indicator from the crystal structure and the colour disappears.</p> <p>Magnesium ion is a fairly strong Brønsted–Lowry acid: <math display="block">[\text{Mg}(\text{H}_2\text{O})_6]^{2+} + \text{H}_2\text{O} \rightleftharpoons [\text{Mg}(\text{H}_2\text{O})_5(\text{OH})]^+ + \text{H}_3\text{O}^+</math> Again, the indicator reacts with the mobile ions but is excluded from the solid crystal lattice.</p>	

## HYDROLYSIS OF tert-BUTYL CHLORIDE

<b>TOPIC</b>	Hydrolysis 2	<b>DEMO #</b> 12.IV.9
<b>REFERENCE</b>	Handout from Dr. Bob Perkins, and A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 210	
<b>EQUIPMENT</b>	500 mL separatory funnel 2–400 mL beaker (one to contain HCl and one to catch liquid from separatory funnel) 1 L beaker to hold ice for cooling HCl distillation apparatus (optional) magnetic stirrer, with stir bar 500 mL graduated cylinder or (600 mL beaker and 250 mL graduated cylinder) 10 mL graduated cylinder dropping pipet	
<b>CHEMICALS</b>	Part 1: 96 mL of tert-butyl chloride 280 mL of concentrated hydrochloric acid ice for cooling 50 mL of 5% sodium hydrogen carbonate (2.5 g NaHCO <sub>3</sub> diluted to 50 mL) 25 g CaCl <sub>2</sub> (anhydrous)  Part 2: 250 mL isopropyl alcohol (70% works well) 1 mL of tert-butyl chloride 4 mL of universal indicator a few millilitres of 1 M NaOH	
<b>PROCEDURE</b>	Part 1: Preparation of tert-Butyl Chloride	

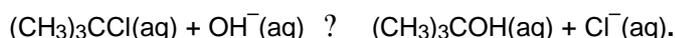
Add 280 mL of cold (5°C) concentrated hydrochloric acid to a separatory funnel and then slowly add 96 mL (1 mol) of tert-butanol with gentle swirling. Allow to stand for 20 minutes, with occasional swirling and occasional relief of pressure. The original one-phase system will gradually separate into 2 layers, with the tert-butyl chloride on top. After 20 minutes, separate the layers (keep the top layer).

Wash the tert-butyl chloride with 50 mL of 5% sodium hydrogen carbonate (CAUTION: foams) and then wash with 50 mL of water. Transfer the tert-butyl chloride to a dry flask and add 25 g of anhydrous calcium chloride to dry the product. Filter the crude tert-butyl chloride. The tert-butyl chloride can be used directly in Part 2 or purified by distillation. Collect the fraction boiling at 49–52°C. The yield should be 40–45 g of colourless liquid.

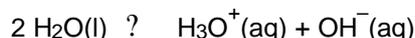
### Part 2: The Hydrolysis Reaction

Into either a 500 mL graduated cylinder or a 600 mL beaker place 250 mL of distilled water and 250 mL of isopropyl alcohol. Place the container on a magnetic stirrer and begin stirring. Add 1 mL of tert-butyl chloride and stir for another 15 seconds, add 4 mL of universal indicator and stir for another 15 seconds. Finally, add a few drops of 1 M NaOH and continue stirring. The initial blue colour will turn to green, orange and finally red. The process can be repeated by adding some additional NaOH to the reaction mixture.

The reaction occurring is



As the hydrolysis proceeds, hydroxide ions are consumed and the solution becomes more acidic as a result of the equilibrium shift in the auto-ionization of water.



## TITRATION OF $\text{Ba}(\text{OH})_2$ WITH $\text{H}_2\text{SO}_4$

<b>TOPIC</b>	Acid-base titrations	<b>DEMO #</b> 12.IV.9
<b>REFERENCE</b>	A Demo A Day: A Year of Chemical Demonstrations, p. 101	
<b>EQUIPMENT</b>	conductivity apparatus magnetic stirrer and spin bar OR stirring rod buret stand and buret clamp 250 mL beaker	
<b>CHEMICALS</b>	100 mL of saturated barium hydroxide, $\text{Ba}(\text{OH})_2$ (about 8 g in 200 mL) 50 mL of 1 M sulphuric acid phenolphthalein solution	
<b>PROCEDURE</b>	Fill the buret with 1 M sulphuric acid. Put 100 mL of saturated barium hydroxide into a 250 mL beaker, add a spin bar and start stirring. Add a few drops of phenolphthalein to the beaker until a persistent colour exists. Lower the conductivity probe into the beaker.  Start adding sulphuric acid from the buret. As the titration progresses, a cloudy precipitate of barium sulphate forms and the conductivity drops. Slow the titration to a dropwise addition until the light goes out and the colour fades. Continued addition of sulphuric acid causes the conductivity to rise as the solution contains an excess of sulphuric acid.	

## GOLDENROD MESSAGES

<b>TOPIC</b>	Indicators 1	<b>DEMO #</b> 12.IV.10
<b>REFERENCE</b>	Twenty Demonstrations Guaranteed to Knock Your Socks Off! , p. 1	
<b>EQUIPMENT</b>	2–3 sheets of goldenrod paper white candle Scotch® tape spray bottle with a “mist” setting	
<b>CHEMICALS</b>	5% ammonia solution	
<b>PROCEDURE</b>	<b>Note:</b> Some brands of goldenrod paper work, some don't. You may have to try several brands.	

Tape 2–3 sheets of goldenrod paper together and flip over the sheets so that the tape cannot be seen. Use a white wax candle to write (rub) a message such as “INDICATORS” on the paper. Brush off any bits of paraffin that could reveal the presence of a message.

Put some 5% ammonia solution in the spray bottle. When students enter, spray the sign with ammonia. The sign becomes red, except where the wax message is written, instantly revealing the message. As the ammonia diffuses out of the solution, the message fades and the sign dries in time for use in the next class.

### ***What is Happening:***

Goldenrod paper contains an indicator dye that changes colour in the presence of basic gases such as ammonia. Exposing the paper to ammonia fumes changes the colour of the paper from yellow–orange to red. If wax is put on the side of the paper exposed to the ammonia vapour, the wax protects the dye from the ammonia and the colour remains unchanged wherever the wax is.

## INDICATOR BOXES

**TOPIC** Indicators 2

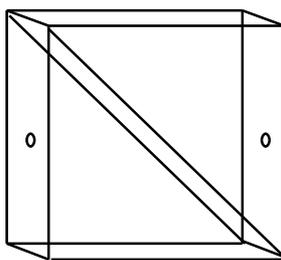
**DEMO #** 12.IV.11

**REFERENCE** commonly known

**EQUIPMENT** Indicator box (see below)

**CHEMICALS** Bromothymol blue (for example)  
0.1 M NaOH  
0.1 M HCl

**PROCEDURE** Construct an indicator box as shown below. The material used is clear plexiglass.



The box is any convenient size, but 25 cm x 25 cm x 5 cm works well. Care must be taken when fitting the diagonal piece inside to prevent leaks. Two holes drilled in opposing sides should accommodate size 00 rubber stoppers.

Fill both triangular portions of the box almost full of distilled water, add about 5-10 drops of the chosen indicator (bromothymol blue works well) to both sides. Then add about 5-10 drops of 0.1M NaOH to one side and 5-10 drops of 0.1 M HCl to the other side. Mix well and finally fill both sides completely full with distilled water.

When viewed from above, one sees the yellow (acidic) colour of bromothymol blue on one side, the blue (basic) colour on the opposite side and the various green colours formed when roughly equal amounts of the acidic and basic forms are present.

## INDICATOR COLOUR CHANGES

<b>TOPIC</b>	Indicators 3	<b>DEMO #</b> 12.IV.12
<b>REFERENCE</b>	commonly known	
<b>EQUIPMENT</b>	4–test tube racks, each able to hold 9 test tubes (18 x 150 mm) 36 test tubes, 18 x 150 mm	
<b>CHEMICALS</b>	pH 3 buffer (250 mL of 0.10 M potassium hydrogen phthalate + 56 mL of 0.20 M HCl / 500 mL) pH 4 buffer (250 mL of 0.10 M potassium hydrogen phthalate + 0.25 mL of 0.20 M HCl / 500 mL) pH 5 buffer (250 mL of 0.10 M potassium hydrogen phthalate + 56 mL of 0.20 M NaOH / 500 mL) pH 6 buffer (250 mL of 0.10 M $\text{KH}_2\text{PO}_4$ + 14 mL of 0.20 M NaOH / 500 mL) pH 7 buffer (250 mL of 0.10 M $\text{KH}_2\text{PO}_4$ + 73 mL of 0.20 M NaOH / 500 mL) pH 8 buffer (250 mL of 0.10 M $\text{KH}_2\text{PO}_4$ + 115 mL of 0.20 M NaOH / 500 mL) pH 9 buffer (250 mL of 0.025 M $\text{Na}_2\text{B}_4\text{O}_7$ + 12 mL of 0.20 M HCl / 500 mL) pH 10 buffer (250 mL of 0.025 M $\text{Na}_2\text{B}_4\text{O}_7$ + 27 mL of 0.20 M NaOH / 500 mL) pH 11 buffer (250 mL of 0.050 M $\text{NaHCO}_3$ + 57 mL of 0.20 M NaOH / 500 mL) methyl orange indicator (0.01% in water) bromothymol blue indicator (0.1 g in 16 mL 0.01 M NaOH + 234 mL water) phenolphthalein indicator (0.05 g in 100 mL of 50% v/v ethanol–water)	
<b>PROCEDURE</b>	Label each of 4 sets of test tubes sequentially: 3, 4, 5, 6, 7, 8, 9, 10, 11.	

Into the 1st test tube set: add a few drops of methyl orange indicator to each test tube.

Into the 2nd test tube set: add a few drops of bromothymol blue indicator to each test tube.

Into the 3rd test tube set: add a few drops of phenolphthalein indicator to each test tube.

Show students how the pH at which an indicator is half–way through its colour change equals its  $\text{p}K_a$  value.

### To show how a “universal indicator” is constructed —

Into the 4th test tube set: add a few drops of methyl orange indicator to each test tube, then a few drops of bromothymol blue and finally a few drops of phenolphthalein.

## PRODUCING AN ACID–BASE INDICATOR PAPER

<b>TOPIC</b>	Indicators 4	<b>DEMO #</b> 12.IV.13
<b>REFERENCE</b>	A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 205	
<b>EQUIPMENT</b>	Mortar and pestle filter paper paper towels scissors sealable container with at least 15 mL capacity, such as a sealable plastic bag	
<b>CHEMICALS</b>	one scoopful of turmeric 5 mL of ethanol (you can also use methanol or isopropyl alcohol) various acids and bases, such as vinegar, oven cleaner, ammonia cleaner, lemon juice	
<b>PROCEDURE</b>	Soak a scoopful of turmeric overnight (or longer) in 5 mL of ethanol, using a sealable container. Grind the turmeric–ethanol mixture in a mortar and pestle to help intensify the colour. Cut the filter paper into strips and thoroughly soak the strips in the turmeric solution. Dry the strips on paper towel. The dry strips are indicator paper.  Test the indicator strips with various acids and bases, including a neutral solution. Turmeric is yellow in acids and red–brown in bases.	

## DILUTING A BUFFER HAS NO EFFECT ON ITS pH

<b>TOPIC</b>	Buffers 1	<b>DEMO #</b> 12.IV.14
<b>REFERENCE</b>	commonly known	
<b>EQUIPMENT</b>	pH meter 2 – 250 mL beakers	
<b>CHEMICALS</b>	pH = 3 (or 4) buffer distilled water	
<b>PROCEDURE</b>	Pour about 25 mL of buffer into a 250 mL beaker. Insert the pH probe and show that the pH agrees with the label on the bottle. Slowly add water: with constant stirring, to show that pH only changes to a very slight extent. Point out that if the total ion concentration were kept constant, the pH would not change at all.	

## COMMON ION EFFECT II

<b>TOPIC</b>	Buffers 2	<b>DEMO #</b> 12.IV.15
<b>REFERENCE</b>	A Demo A Day: A Year of Chemical Demonstrations, p. 214	
<b>EQUIPMENT</b>	3 – 100 mL graduated cylinders 3 – 250 mL graduated cylinders dropping pipet	
<b>CHEMICALS</b>	2 mL of universal indicator 75 g of calcium carbonate, $\text{CaCO}_3$ 180 mL of 2 M acetic acid, $\text{CH}_3\text{COOH}$ 27.0 g of sodium acetate trihydrate, $\text{NaCH}_3\text{COO}\cdot 3\text{H}_2\text{O}$ 100 mL of 2 M hydrochloric acid 2 mL of universal indicator 20 mL of distilled water	
<b>PROCEDURE</b>	<p>Mix 27.0 g of <math>\text{NaCH}_3\text{COO}\cdot 3\text{H}_2\text{O}</math> and 80 mL of 2 M <math>\text{CH}_3\text{COOH}</math> and dilute to 100 mL with water. Pour this mixture into the 1st 100 mL graduated cylinder. To the 2nd 100 mL graduated cylinder add 100 mL of 2 M HCl. To the 3rd 100 mL graduated cylinder add 100 mL of 2 M <math>\text{CH}_3\text{COOH}</math>.</p> <p>To each 250 mL graduated cylinder add 10 drops of universal indicator and add 25 g of powdered <math>\text{CaCO}_3</math>.</p> <p>Have three students simultaneously pour the contents of each 100 mL graduated cylinder into a different 250 mL graduated cylinder. The HCl froths instantly, the acetic acid froths somewhat slower and the acetic acid–sodium acetate buffer froths significantly slower.</p>	

## **$K_{SP}$ BALANCE REVISITED**

<b>TOPIC</b>	Buffers 3	<b>DEMO #</b> 12.IV.16
<b>REFERENCE</b>	A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 236	
<b>EQUIPMENT</b>	2 – 250 mL beakers 2 – 100 mL graduated cylinders 1 – 25 mL graduated cylinder	
<b>CHEMICALS</b>	0.50 M ammonium chloride: $\text{NH}_4\text{Cl}$ (6.69 g $\text{NH}_4\text{Cl}$ / 250 mL) 1.0 M ammonium hydroxide: $\text{NH}_4\text{OH}$ (33.8 mL conc. $\text{NH}_4\text{OH}$ / 250 mL) 0.50 M magnesium sulphate: $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ (30.8 g $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ / 250 mL)	
<b>PROCEDURE</b>	Put 20 mL 0.5 M $\text{MgSO}_4$ in beaker and add 20 mL 1 M $\text{NH}_4\text{OH}$ . Precipitate forms because of available $\text{OH}^-$ ions.  In second beaker place 20 mL 1 M $\text{NH}_4\text{OH}$ and 20 mL 0.5 M $\text{NH}_4\text{Cl}$ . Add 20 mL 0.5 M $\text{MgSO}_4$ and observe that no precipitate forms because the buffer has too few $\text{OH}^-$ ions.	

## THE EFFECT OF A CHANGE IN [CO<sub>2</sub>] ON BLOOD pH

**TOPIC** Buffers in biological systems **DEMO #** 12.IV.17

**REFERENCE** Chemical Demonstrations: A Sourcebook for Teachers, Volume 1, p. 88

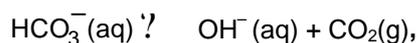
**EQUIPMENT** 500 mL side-arm vacuum filtration flask  
rubber stopper to fit flask  
rubber tubing to fit flask to aspirator pump

**CHEMICALS** 150 mL of saturated sodium hydrogen carbonate (about 24 g of NaHCO<sub>3</sub> diluted to 150 mL)  
phenolphthalein solution  
1 M hydrochloric acid

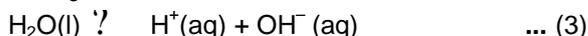
**PROCEDURE** Place 150 mL of saturated sodium hydrogen carbonate in the vacuum flask and connect the flask to the aspirator pump. Add 3–4 drops of phenolphthalein solution to the flask. If the solution is pink, add just enough 1 M HCl to just turn the solution colourless.

Stopper the flask, turn on the aspirator pump and notice that gas bubbles are formed and the solution turns pink, indicating the solution is now more basic.

The overall equilibrium is



which is the result of the following three interlinked equilibria.



The loss of gaseous carbon dioxide shifts equilibrium (1) to the products side, decreasing the concentration of dissolved carbon dioxide, which in turn shifts equilibrium (2) to the product side and decreases the concentration of hydronium ion. As a result of the loss of hydronium ion, equilibrium (3) shifts to the products side, increases the concentration of hydroxide ion, increases the pH and turns the phenolphthalein pink.

## BASIC ANHYDRIDE

<b>TOPIC</b>	Basic anhydride	<b>DEMO #</b> 12.IV.18
<b>REFERENCE</b>	A Demo A Day – A Year of Physical Science Demonstrations, p. 188	
<b>EQUIPMENT</b>	stand with ring and a clamp bunsen burner and flint striker small pyrex funnel wash bottle with distilled water 2 – 15 x 150 mm test tubes, one with a stopper to fit	
<b>CHEMICALS</b>	magnesium ribbon, 10 cm universal indicator, 1 mL	
<b>PROCEDURE</b>	<p>Clamp the test tube to the stand at the bottom. Place the funnel into the test tube. Clamp the ring about 10 cm above the top of the funnel. Wrap the end of a 10 cm piece of magnesium ribbon around the ring, in such a way that the free end of the ribbon is centred over the funnel.</p> <p>Use the bunsen burner to ignite the magnesium, warning students not to look directly at the light. Wash the ashes into the test tube and fill the tube half–full of distilled water. Stopper the tube and shake the mixture vigorously.</p> <p>Put some distilled water in a second test tube and add universal indicator to both the distilled water and the mixture of magnesium oxide and water. The magnesium oxide solution is basic.</p>	

## ACID RAIN

<b>TOPIC</b>	Acid rain 1	<b>DEMO #</b> 12.IV.19
<b>REFERENCE</b>	A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 216	
<b>EQUIPMENT</b>	15 x 100 mm test tube 2-hole stopper to fit test tube 500 mL erlenmeyer flask 1-hole stopper to fit flask 10 cm length of glass tubing to fit through the one hole stopper in the flask and through one hole of the two-hole stopper. Large plastic dropper	
<b>CHEMICALS</b>	1 g of copper turnings or wire 1 mL of bromcresol green indicator solution 1 mL of 1 M ammonium hydroxide 1 mL of concentrated nitric acid glycerine	
<b>PROCEDURE</b>	<p>When the apparatus is assembled, the test tube is inside the flask. The test tube has a 2-hole stopper in it, a piece of glass tubing goes down through the stopper just inside the test tube, the upper end of the glass tubing goes up through a 1-hole stopper in the flask and the glass tubing extends about 4–5 cm above the one-hole stopper. Place 1 g of copper inside the test tube and stopper it with the 2-hole stopper (with the glass tube and one-hole stopper attached). Pour 100 mL of water into the flask and add 1 mL of bromcresol green indicator. Add just enough 1 M ammonium hydroxide solution to just turn the solution blue. Place the test tube assembly into the flask and seal the flask with the one-hole stopper. Fill a plastic dropper with 1 mL of concentrated acid and place the dropper into the glass tubing extending up from the stopper in the flask.</p> <p>Slowly drip the nitric acid onto the copper turnings. Avoid a vigorous reaction. Red-brown nitrogen dioxide is evolved and escapes into the flask. When the flask is swirled, the gas dissolves into the water and turns the solution yellow, indicating the presence of an acid.</p>	

## INSTANT SMOG

**TOPIC** Acid rain 2

**DEMO #** 12.IV.20

**REFERENCE** A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 216

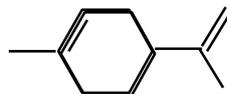
**EQUIPMENT** tesla coil  
aluminum foil  
500 or 100 mL erlenmeyer flask, with a rubber stopper to fit  
copper wire

**CHEMICALS** fresh lemon rind  
source of oxygen gas (optional)

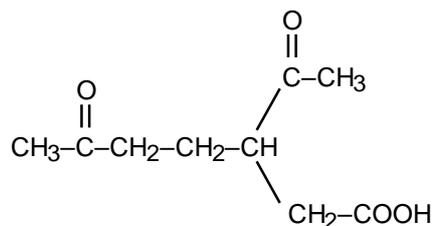
**PROCEDURE** Wrap the outside of the flask with aluminum foil and wrap a piece of copper wire around the neck of the flask, to ground the flask. Connect the wire to a water tap. Make a coil out of copper wire and put the copper coil into the flask, such that the wire protrudes from the top. Spark the copper coil with a tesla coil for either one minute (if the flask is filled ahead of time with oxygen) or three minutes (if ambient air is used). Remove the wire and quickly stopper the flask. Remove the wire and aluminum foil from the outside of the flask. The flask now contains a high percentage of ozone.

Use a knife to peel a twist of lemon rind. Remove the stopper from the flask, sharply twist or crush the lemon peel, drop the peel into the flask and restopper the flask. Within seconds a wispy white cloud forms inside the flask, and eventually creates a haze in the flask.

The haze is likely due to ozonolysis of the double bonds in the limonene in the lemon rind. The structure of limonene is shown below.



The oxidation by ozone produces 3-acyl-6-one-heptanoic acid, shown below.



As a result of the low vapour pressure of this and other possible products, the molecules spontaneously form a visible aerosol as nucleation occurs around particles in the atmosphere.

## THE SILVER NITRATE / COPPER REACTION REVISITED

**TOPIC** Redox reactions

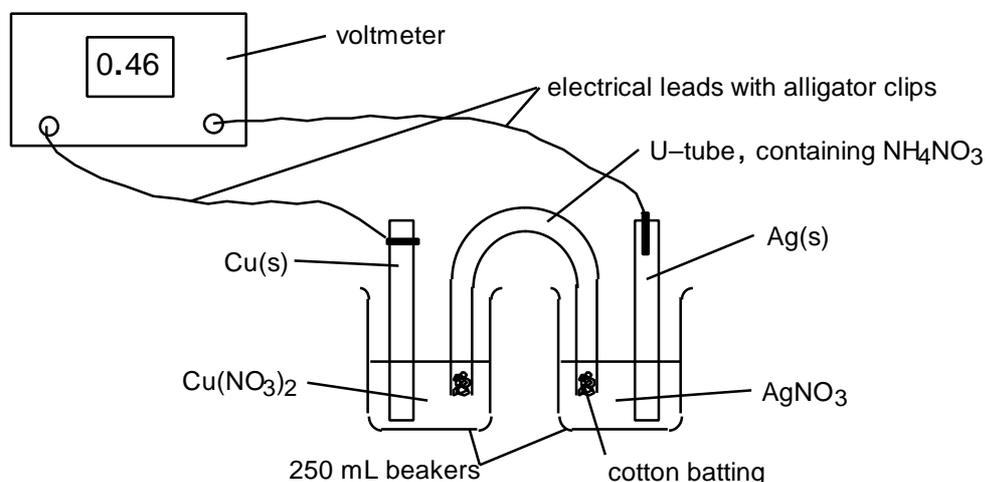
**DEMO #** 12.V.1

**REFERENCE** commonly known

**EQUIPMENT** 2–250 mL beakers  
U-tube  
voltmeter, with 2 electrical leads ending in alligator clips  
cotton batting

**CHEMICALS** piece of copper sheet (about 2 cm x 12 cm) or 12 cm length of very thick wire  
piece of silver sheet (about 2 cm x 12 cm) or 12 cm length of silver wire  
250 mL of 0.5 M ammonium nitrate ( $\text{NH}_4\text{NO}_3$ ) solution  
100 mL of 0.5 M silver nitrate ( $\text{AgNO}_3$ ) solution (a 0.1 M solution can also be used)  
100 mL of 0.5 M copper (II) nitrate ( $\text{Cu}(\text{NO}_3)_2$ ) solution

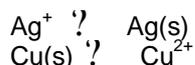
**PROCEDURE** Set up the demonstration as shown below.



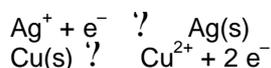
From Chem 11, students should be aware of the single replacement reaction:



Write the following incomplete half-reactions on the blackboard or overhead:



and ask students to finish balancing the half-reactions. Once the half-reactions are balanced



ask students which way the electrons are flowing through each of the two wires (from the copper to the silver). Point out that one electron is involved in the silver half-reaction while two electrons are involved in the copper half-reaction. Ask how this problem can be resolved (double the silver half-reaction).

Finally, ask whether the copper metal is necessary for this reaction to occur (yes). Is the  $\text{Cu}^{2+}$  solution necessary? (No, any electrolyte can be used. To prove that the  $\text{Cu}^{2+}$  is not required, replace the  $\text{Cu}(\text{NO}_3)_2$  solution with  $\text{NH}_4\text{NO}_3$  solution. The voltage might be slightly different but the same reaction continues.) Is the  $\text{AgNO}_3$  solution necessary? (Yes, because it supplies the  $\text{Ag}^+$  ions used in the reaction.) Is the silver metal necessary? (No, any electrical conductor will do. To show this, remove the piece of silver and put the alligator clip directly into the solution.)

## REDUCED IRON IN CEREAL

**TOPIC** Reduced iron **DEMO #** 12.V.2

**REFERENCE** commonly known

**EQUIPMENT** magnetic stirrer with teflon coated stirring bar  
mortar and pestle  
2 L beaker

**CHEMICALS** about 250 mL of Cheerios™ or other cereal containing “reduced iron” according to its labelled contents

**PROCEDURE** Finely grind about 250 mL of cereal in a mortar and pestle and add it to a 2 L beaker containing about 1500 mL of water and a stirring bar. Stir the mixture for about one minute and retrieve the stirring bar. There will be fine black particles on the bar. This black powder is metallic iron.

**Explanation:** The iron powder (“reduced iron”) is put into cereals to react with the hydrochloric acid in our stomach and provide  $\text{Fe}^{2+}$ , a biological nutrient. Iron (II) cannot be put in the cereal in chemical form because  $\text{Fe}^{2+}$  is rapidly oxidized to  $\text{Fe}^{3+}$ , which is unusable as a nutrient.

## THE MANY COLOURS OF VANADIUM

<b>TOPIC</b>	Oxidation numbers	<b>DEMO #</b> 12.V.3
<b>REFERENCE</b>	Chemical Curiosities: Spectacular Experiments and Inspired Quotes, p. 62	
<b>EQUIPMENT</b>	250 mL graduated cylinder 10 mL graduated cylinder long stirring rod	
<b>CHEMICALS</b>	1.0 g of ammonium metavanadate, $\text{NH}_4\text{VO}_3$ 2 mL concentrated sulphuric acid 100 mL of 6 M hydrochloric acid zinc granules	
<b>PROCEDURE</b>	Dissolve or suspend 1.0 g of ammonium metavanadate in 200 mL of distilled water. (The solution will be colourless initially.) Stir in 2.0 mL of concentrated sulphuric acid. (The solution turns yellow and a small amount of red precipitate forms.) Add 6–8 zinc granules, followed by 25 mL of 6 M hydrochloric acid, followed by a further 40–50 mL of 6 M hydrochloric acid. (Rapid gas evolution occurs and the mixture turns sky-blue, then green and finally turquoise followed by lilac for a short time before a return to turquoise.	

$\text{VO}_3^?$  = vanadium (VII) = colourless

$\text{VO}_2^?$  = vanadium (V) = yellow

$\text{V}_2\text{O}_5$  = red precipitate

$[\text{VO}(\text{H}_2\text{O})_5]^{2+}$  = vanadium (IV) = sky-blue

$\text{V}^{3+}(\text{aq})$  = green

$\text{V}^{2+}(\text{aq})$  = lilac

## THE BREATHALYSER

<b>TOPIC</b>	Applied Electrochemistry	<b>DEMO #</b> 12.V.4
<b>REFERENCE</b>	Chem 13 News, January 1979, p. 15	
<b>EQUIPMENT</b>	Test tube rack with 6–18 x 150 mm test tubes	
<b>CHEMICALS</b>	250 mL of breathalyser solution, made ahead of time. If sealed very well, the solution can be stored for over a year. The solution is identical to the solution formerly used by law enforcement officers and is made as follows.	

Dissolve 6.25 g of potassium dichromate in 125 mL of distilled water. Slowly and carefully add 125 mL of concentrated sulphuric acid, small portions at a time, TO the above solution, with constant cooling under cold running water. (CARE: The solution gets very hot and is extremely corrosive!) After the addition of acid is finished and the solution has cooled, add 10 mL of silver nitrate solution (2 g of silver nitrate dissolved in 10 mL of water).

**PROCEDURE** To duplicate the police procedure, the reaction must take place at  $50 \pm 3^\circ\text{C}$ , but the reaction actually works quite well at room temperature.

To each test tube add 3-5 mL of breathalyser solution and then add a few drops of each compound to be tested.

If the compound is an alcohol, the yellow breathalyser solution is rapidly oxidized to a green colour. The quicker the colour change and the darker the final green colour, the more alcohol is present. Substances such as acetone or ethyl acetate will eventually cause a colour change but they are much slower to react and can easily be distinguished from alcohols.

Some solutions to try:

- cirrhosis)
- ethanol (beverage alcohol) — reacts very quickly (toxic to the liver in large quantities —
  - methanol (fondue fuel) — reacts very quickly (very toxic, ingestion causes blindness)
  - 2-propanone (acetone) — does not react for a long time
  - 2-propanol (isopropyl alcohol, i.e. rubbing alcohol) — reacts very quickly (very toxic)
  - ethyl acetate (i.e. nail polish remover) — reacts very slowly (very toxic)
  - 2-butanone (methyl ethyl ketone) — reacts very slowly (very toxic)

**NOTE:** A person who is shown to be impaired by the breathalyser test theoretically could have consumed any alcohol, but all alcohols except for ethanol are immediately toxic so that only ethanol will allow a person to be both alcohol impaired and alive.

Incidentally, if a person has consumed methanol the medical treatment involves administering massive and continuing amounts of ethanol, which keep the person very drunk, accompanied by monitoring the urine output for excretion of methanol. Oxidation of methanol by the liver creates methanal ("formaldehyde"), which is extremely toxic. If large amounts of ethanol are also present in the system in a continuous manner, the liver is preoccupied with oxidizing the ethanol to ethanal ("acetaldehyde"), which is much less toxic than methanal, and the methanol is simply excreted in unoxidized form. Once the urine coming from the catheter in the victim shows that the last of the methanol has been excreted, the victim is sobered up.

## HOW TO TURN ALUMINUM INTO HOARFROST

<b>TOPIC</b>	Corrosion	<b>DEMO #</b> 12.V.5
<b>REFERENCE</b>	Chemical Curiosities: Spectacular Experiments and Inspired Quotes, p. 232	
<b>EQUIPMENT</b>	A small sheet of aluminum (not aluminum foil), about 10–15 cm square dropper emery cloth	
<b>CHEMICALS</b> water)	50 mL of saturated mercury (II) chloride, $\text{HgCl}_2$ (about 3–4 g dissolved in 50 mL of 2 mL of 1 M hydrochloric acid	
<b>PROCEDURE</b>	<p>Polish one side of an aluminum sheet with emery cloth. Mix 2 mL of 1 M hydrochloric acid into 50 mL of saturated mercuric chloride. [Note: If tightly stoppered, this solution can be used for many years, although a bit more HCl may need to be added if the desired reaction does not occur in subsequent years.] Use a dropper to place several drops of the acidic mercury solution on the surface of the aluminum plate. The aluminum almost immediately takes on a dull, dark grey colour. After about 10 minutes a frostlike grey–white growth about 2-5 mm high appears where the drops were added and needles begin to grow.</p> <p>The acid destroys the protective oxide layer on the aluminum and galvanic action allows the aluminum to be oxidized while <math>\text{Hg}^{2+}</math> is reduced to grey metallic mercury. The aluminum is converted to <math>[\text{Al}(\text{H}_2\text{O})_6]\text{Cl}_3</math>, which is responsible for the appearance of the white needles.</p>	

## ELECTROLYSIS OF WATER

**TOPIC** Electrolysis 1 **DEMO #** 12.V.6

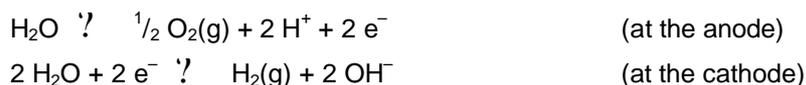
**REFERENCE** commonly known

**EQUIPMENT** U-tube  
power supply, with 2 electrical leads ending in alligator clips  
2-carbon rods  
stand with clamp to secure U-tube  
large beaker (big enough to hold U-tube, carbon rods and solution)

**CHEMICALS** bromothymol blue indicator  
250 mL of 0.5 M sodium sulphate (17.8 g Na<sub>2</sub>SO<sub>4</sub> / 250 mL)

**PROCEDURE** Clamp the U-tube about half-way down the tube. Add several drops of bromothymol blue to the sodium sulphate solution until a definite green colour is present. Pour the sodium sulphate solution into the U-tube until it is almost full (about 1 cm from the top). Use the alligator clips to hold the carbon rods at a depth of about 3-4 cm into the solution in each arm of the U-tube. Connect the power supply and pass about 3-5 volts through the solution for about 2 to 3 minutes.

The half-reactions occurring are shown below.



Point out that bromothymol blue is yellow in acidic solutions and blue in basic solutions. Ask students which side of the U-tube is the anode (yellow side) and which is the cathode (blue side).

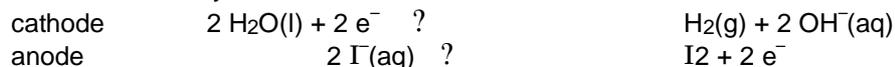
Ask students to predict what colour will form if the contents of the U-tube are poured into a beaker and mixed. (green) Pour the contents of the U-tube into a large beaker so as to avoid spilling even a drop of liquid. Make sure the carbon rods also go into the beaker because the carbon may absorb some of the acid or base and there may be an acid/base deficit if the rods are not included. As the last drops come out, the solution reverts to its previous green colour, confirming student predictions.

## WRITING WITH A HALOGEN

- TOPIC** Electrolysis 2 DEMO # 12.V.7
- REFERENCE** A Demo A Day, Volume 2: Another Year of Chemical Demonstrations, p. 266
- EQUIPMENT** HB pencil, dull, which has been cut with a knife toward the top or middle, in such a way that an alligator clip can make electrical contact with the graphite  
piece of aluminum foil, about 15 cm square  
9 V battery  
2 – electrical leads with alligator clips on both ends  
filter paper (about 12-15 cm)
- CHEMICALS** 50 mL of 0.1 M KI solution (0.8 g diluted to 50 mL)  
100 mL of 1 % starch solution (Add 1 g starch to 5 mL of water (to make a paste) and  
add  
to 75 mL of boiling water; boil for 2 minutes, cool and dilute to 100 mL)
- PROCEDURE** Mix 3 mL of the starch solution into the KI solution. Lay the filter paper on top of the aluminum foil and add several drops of the starch–KI solution to the filter paper so as to wet the filter paper. Connect the negative battery terminal to the aluminum foil with an electrical lead and connect the positive terminal to the lead pencil with the second electrical lead.

Using the pencil as a stylus, lightly touch the paper and write a message, which will show as a purple–black line. Show students that the same pressure on the dry part of the filter paper does not a noticeable pencil mark.

The equations for the electrolysis reaction are



and the iodine produced reacts with the starch to produce the well known starch–iodine colour.

## ELECTROLYSIS OF COPPER (II) CHROMATE

<b>TOPIC</b>	Electrolysis 2	<b>DEMO #</b> 12.V.8
<b>REFERENCE</b>	Journal of Chemical Education, Vol. 78. No. 2, February 2001, p. 207	
<b>EQUIPMENT</b>	2–Petri dishes 2–opened paper clips filter paper: 2 cm wide and 15 cm long glass capillary DC power supply, capable of delivering 30 V, with two electrical leads ending in alligator clips	
<b>CHEMICALS</b>	copper (II) chromate: Made by mixing 25 mL of 1 M CuSO <sub>4</sub> and 25 mL of 1 M K <sub>2</sub> CrO <sub>4</sub> . Filter and wash the red–brown precipitate of CuCrO <sub>4</sub> . To do the demonstration, put a small amount of the solid in a vial or small test tube and dissolve in a few drops of concentrated ammonia to produce a dark green solution. Electrolysis electrolyte: Dissolve 2 g of NH <sub>4</sub> Cl in 50 mL of water and 5 mL of concentrated NH <sub>3</sub> .	
<b>PROCEDURE</b>	Draw a line across the middle of the filter paper strip, immerse in the electrolyte and put on a piece of paper towel to get rid of excess liquid. Place two Petri plates 3 cm apart and fill with electrolyte. Open two paper clips and place in opposite ends of the Petri plates to act as electrodes, such that most of the clips are in the solution and only a little remains out of the solution to be attached to alligator clips. Use the capillary to draw up some of the ammoniacal CuCrO <sub>4</sub> solution and use it to draw a line over the pencil mark with the solution in the capillary. Use the filter paper strip as a salt bridge between the two Petri plates, immersing the ends in the solution in the plates and aligning the pencil mark midway between the plates. Connect the alligator clips to the paper clips and power supply, and then turn on the power supply (using 30–50 V; CARE!).  After 1-3 minutes, yellow (CrO <sub>4</sub> <sup>2-</sup> ) and blue (Cu(NH <sub>3</sub> ) <sub>4</sub> <sup>2+</sup> ) bands can be seen migrating away from each other. If the polarity of the current is reversed, the ions will move toward each other, form a green band and then pass through each other as they again separate in the opposite direction.  This demonstration is an excellent way to visibly prove that anions and cations travel in opposite directions through a salt bridge.	

## ELECTROLYTIC AND ELECTROCHEMICAL CELLS

**TOPIC** Electrolytic vs electrochemical cells **DEMO #** 12.V.9

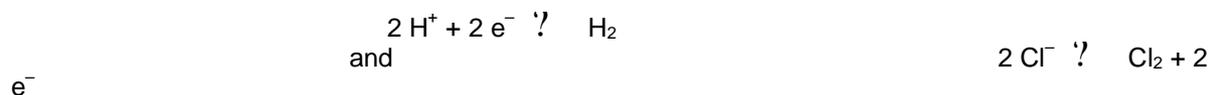
**REFERENCE** A Demo A Day: A Year of Chemical Demonstrations, p. 248

**EQUIPMENT** voltmeter  
 3 to 5 V DC power supply  
 2 – connection wires with alligator clips  
 2 – porous carbon rods  
 600 mL beaker

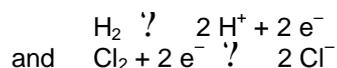
**CHEMICALS** 400 mL of 6 M hydrochloric acid

**PROCEDURE** Pour 400 mL of 6 M HCl into a 600 mL beaker. Place porous carbon rods as deep into the acid solution and as far apart as possible. Connect the DC power supply to the carbon rods with the alligator clips and electrolyze the solution for 3–4 minutes. Next, remove the power supply and replace it with a voltmeter. The system now operates as a galvanic cell and reads close to 1.36 V.

As the electrolysis proceeds, the half-cell reactions are as shown below.



The gases produced are absorbed into the porous carbon rods. After the power supply is removed, the rods act as their own half cell in a galvanic reaction involving the spontaneous oxidation of hydrogen and reduction of chlorine.



## PETRI DISH ELECTROLYSIS and LIQUID MOTORS

<b>TOPIC</b>	Electrolysis 3	<b>DEMO #</b> 12.V.10
<b>REFERENCE</b>	Twenty Demonstrations Guaranteed to Knock Your Socks Off! Volume II, p. 21	
<b>EQUIPMENT</b>	9 V battery clip for 9 V battery small plastic petri plate 2 small paper clips neodymium magnet two 30–40 cm lengths of copper wire, bared at both ends (or better, two lengths of wire having alligator clips at both ends)	
<b>CHEMICALS</b>	250 mL of saturated tin (II) chloride dihydrate, $\text{SnCl}_2 \cdot 2\text{H}_2\text{O}$ . Make freshly by dissolving about 6.5g of powdered $\text{SnCl}_2 \cdot 2\text{H}_2\text{O}$ in 250 mL of distilled water, adding about 0.3 mL of concentrated nitric acid, shaking for about a minute and filtering the resulting suspension to get a clear solution. 250 mL of 0.25 M sodium sulphate	
<b>PROCEDURE</b>	<p>Attach a paper clip to opposite sides of the petri plate, such that the long side of the clip is inside the petri plate and as close to the bottom as possible. Next attach each side of the battery to a paper clip, using wires or alligator clips.</p> <p>Pour 15–20 mL of saturated tin (II) chloride solution into the petri plate. Almost immediately, beautiful fern-like crystals of lustrous metallic tin form at the negative pole and branch their way across the petri plate. At the positive pole a milky-white precipitate forms. The precipitate is probably a mixture of tin (IV) chloride and other tin (IV) compounds.</p> <p><b>Variations:</b></p> <ol style="list-style-type: none"><li>1. Repeat, using a 2 cm length of bare copper wire in the centre of the solution, pointing to the two paper clip electrodes. The two ends of the wire will act as anode and cathode, indicating an electrical flow exists in the wire (which has less resistance than the solution).</li><li>2. Allow the electrolysis to continue for about 30–40 seconds and then reverse the polarity of the paper clips by switching the connections to the battery. The tin crystals will be seen to corrode and go into solution, while new crystals appear at the other electrode: electrorefining.</li><li>3. The reaction is incredibly beautiful when seen under a dissecting microscope.</li><li>4. Try electrolyzing sodium sulphate solution having some bromthymol blue. After a while, the anode becomes yellow (acidic) and the cathode becomes blue (basic). Turn off the electricity and stir the solution to get a neutral solution again: the amount of acid generated equals the amount of base generated.</li><li>5. <b>LIQUID MOTORS.</b> This demonstration works wonderfully on the overhead projector. Try running the electrolysis of sodium sulphate centred over a powerful neodymium magnet. As the ions accelerate toward the electrodes, the magnetic field deflects them in a perpendicular manner and causes the entire solution to start swirling (depending on the direction of the current and the orientation of the magnet). Reversing the polarity of the paper clips causes the vortices to reverse their directions. With the magnet aligned toward the two electrodes, a circular motion is seen; with the magnet aligned perpendicular to the electrodes, four vortices are observed; with the magnet oriented vertically, a double vortex is observed. Sprinkling a bit of powdered sulphur on the surface permits the liquid flow to be seen clearly.</li></ol>	