The Farewell Demos

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Edit any lab or demonstration you may wish to do for your purposes. For example take out answers so the students can think about them.
Try each before having students do them or before you demonstrate them to make sure everything will go as expected.
Also remember that any demonstration can be reworked into a hands on lab for students and any lab for students can be demonstrated by the teacher.
Any questions about these labs and demos should be directed towards myself or another experienced chemistry teacher that you have contacts with.

Demo or Lab: CORROSION LAB

Problem: What are the factors that slow and speed corrosion?

Materials: eight clean test tubes, distilled water, tap water, common nails cleaned to remove any oil coating, painted nails, galvanized nails, copper wire, salt, acid, magnesium ribbon, 0.10 mol/L potassium ferricyanide solution

Safety: Follow general safety procedures.

Procedure: Prelab all the set up so students can come in the next day and start the set up.

1. Place a common nail in tap water in a test tube, with the nail head just out of the water
2. Place a 2nd nail in a second test tube completely covered by the tap water
3. Place a 3rd nail in distilled water in a test tube, with the nail head just out of the water
4. Place a 4th nail in the test tube, completely in tap water with salt added
5. Place a 5th nail in the test tube, completely in tap water with 5 drops of acid [vinegar will do]
6. Place a 6th nail wrapped in clean bare copper wire, completely in tap water
7. Place a 7th painted nail completely in salt water
8. Place a 8th galvanized (zinc coated) nail completely in salt water
9. With 10 minutes left in the period, add five drops of the potassium ferricyanide solution to each test tube. The blue colour indicates the presence of rust.

Observations:

1. Make a chart with the test tube number, the appearance before and after the potassium ferricyanide solution was added.
2. Note how much blue and thus how much rusting took place in each test tube.

POST LAB: Review the observations and discuss the answers to the questions, Are there any other thoughts and questions that the students have after performing this activity.

Questions:

1. List the nails in the test tubes in order from the most rusted to least rusted based on the amount of blue material present after the potassium ferricyanide solution was added.
   Acid, salt, contact with a less active metal [copper in this case] promote rusting. Galvanizing, painting, contact with a more active metal [magnesium in this case] decreases corrosion. (see also 32 and #3.)
2. What factors make iron corrode more?
3. What factors protect iron from corrosion?
4. What does acid rain do to steel (iron) bridges and cars?
   The acid promotes rusting thereby decreasing the life of the cars and bridges ans weakens their structure.
5. What are cars and bridges painted?
   Painting protects the cars and bridges from rusting and lengthens their life by preventing moisture and oxygen from coming in contact with the iron.
6. Why are pieces of zinc metal attached to steel ships?
   The ships are made of iron and spend their time in water, which promotes rusting. Many ships are in salt water, which promotes rusting even more. The contact with the zinc provides anodic protection whereby the more active zinc rusts instead of the iron hull of the ship. This is also done with oil and gas pipelines, water lines in many areas.
7. Why are new nails and iron objects often oiled?
   The oil provides a barrier between the air and the metal keeping moisture and oxygen away from the iron to slow corrosion.
8. Why is salt put on roads in the winter? What problem with metal can the salt cause?
   Salt promotes the melting of snow and ice, which helps improve contact with the road in winter. More friction, more control and safer driving. The salt and water however promotes rusting of iron vehicles and bridges.

You can expand on this lab by using a nail in a test tube of air, a nail with oil or Vaseline on it, a bent nail [strain promotes rusting], another road deicer such as potassium acetate, calcium chloride, or any other idea from yourself a college or even better the students.

Demo: Visual Dissolving of a Copper Anode:

Purpose: to demonstrate the dissolving of a copper anode.

Copper is purified by dissolving the copper off an impure copper anode in a copper(II) sulfate solution and depositing it onto a copper cathode as pure copper. There are labs that ask students to set up an electrolytic cell with copper(II) sulfate solution and two copper electrodes. The dissolution of the copper at the anode is difficult to observe. Benedict's solution is a basic solution of a very small amount of copper(II) sulfate in sodium citrate solution. The copper(II) citrate complex is a more intense blue than the hydrated copper(II) ion. If you do not have sodium citrate, potassium sodium tartrate solution can also be used as the copper(II) tartrate complex is also very intense in colour [Fehling's solution]. Ammonia solution could be used since the tetramine copper(II) complex is very dark [Mediteranian blue] but is not recommended because of the strong odour.

Materials: 2 copper electrodes, 0.20 mol/L sodium citrate solution, power supply, U-tube, satnd, clamp, electric connectors

Safety: Wear goggles, do not touch metal parts when power is on

Method: 1. Set the U-tube up on the stand.
2. Add sodium citrate solution to the U-tube until 2.5 cm from the openings.
3. Place the copper electrodes into the U-tube.
4. Attach the electric connectors to the power supply and the copper electrodes.
5. Turn on the power.
6. Observe the reactions at the anode and the cathode.
7. Record your observations.
8. Turn off the power supply. Disconnect the wires, Clean and return the materials.

Observations: The anode will give a blue colour that increases and diffuses through the solution. The cathode gives off a gas.

Questions: Explain the results.

The anode has the copper giving up electrons [oxidation]

$$\mathbf{Cu\ (s) \rightarrow Cu^{+2}\ (aq) + 2\ e^{-}}$$

While at the cathode, water picks up electrons to produce hydrogen gas:

$$2\ H_2O\ +\ 2\ e^-\ \rightarrow\ H_2\ +\ 2\ OH^-$$

The copper- citrate complex shows blue colour of dissolved copper ions.
If left long enough, the solution will become all blue and copper will start to plate out at the cathode.

If the solution is made basic with sodium carbonate solid, you have some Benedict's solution for reducing sugar tests.

Demo: Why Do Ionic Solutions Conduct Electricity?

Purpose: to electrolyse a copper(ii) Chloride solution and explain the conduction of electricity through ionic solutions and liquids.

Materials: carbon electrodes, distilled water, U- tube, stand, clamp, power supply, electric leads, copper(ii) chloride, balance, scoopula, beaker

Safety: wear goggles, be careful of the power supply, do not get shocked.

Method: Prepare 100 mL of a 0.5 mol/L copper(ii) chloride solution. [How many moles? What mass per mole? What mass in 100 mL?] *[.05 mole, 170.5 g/ mole if* $CuCl_2.2H_2O$*, 8.53 g]*
Set up the U – tube on the ground supported on the stand and add the copper(ii) chloride solution until 2.5 cm from the top. Place the left over copper(ii) chloride solution in the bottle provided.
Set the electrodes into the solution. Attach the wires and turn on the power supply. Let stand for 30 minutes with the power running.
Turn the power off. Unplug the power supply, disconnect the wires, take out the electrodes. Note the deposit and odour of the electrodes.

Observations: Note and record the colour of the deposit and odour of the electrodes.

Questions: 1. What is the deposit?
The deposit is copper metal/
2. Where did the deposit come from? What must have happened to form the deposit? Write a chemical equation known as a half reaction, for what took place on this electrode. Was the electrode positive or negative? Explain.
The copper ions in the solution were attracted to the negative electrode [Cathode] and picked up two electrons producing the copper metal. $Cu^{+2}\ +\ 2\ e^-\ \rightarrow\ Cu^0$
3. What was the odour on the other electrode? What material produces this odour? How did this material form? Write a chemical equation for the reaction which occurred. Was this electrode positive or negative?
The odour was that of chlorine. The negative chloride ions migrated to the positive electrode where they gave up electrons to form chlorine molecules. $2\ Cl^-\ \rightarrow\ Cl_2\ +\ 2\ e^-$.

4. Write a chemical equation for the total reaction which occurred. [The sum of the two half reaction equations.]

$$CuCl_2 \rightarrow Cu + Cl_2$$

5. What happens to an electrical circuit if it is broken? Why?
The circuit stops working because the electrons can no longer flow in a complete path.
6. There was a large gap in the electrical circuit in the U – tube. Explain with the use of a diagram how the circuit is made complete because of the chemical reactions which occurred.
The electrons complete the circuit by being picked up from the chloride ions at the anode and by being passed on to the copper(II) ions at the cathode.
7. There are salespeople who go door to door selling water purification systems to people who show the potential customer “hidden impurities” in their water by hooking up iron electrodes to the person’s tap water. After running a current through the system, brown “impurities” appear. What must the brown material be and how is it formed? Is the water really full of “hidden Impurities”? [You may set this electrolysis up in a U – tube if you wish to observe it.]
The iron nail begins to dissolve at the anode and forms rust which clouds the water with a brown solid. The water does not have hidden impurities.

Demo: Acid or Base, Which is Safer?

This demo is done to help students understand that bases can be just as dangerous as acids. They are actually more dangerous because we have more nerve receptors to tell us an acid is causing damage than base receptors. As such we have some resistance to acids but none to bases. Thus an acid spill in the eye may require 15 minutes of washing, a base spill will require 20 to 30 minutes of washing.
Students have the idea that since bases neutralize acids, bases are safer. Not so. Either extreme of the pH scale is dangerous.

Safety: 3.0 mol/L acid and base used so wear goggles, use gloves, use a well ventilated area since acid and base fumes will be sent into the air.

Hydrochloric acid must be used since the chloride ions help break up the protective oxide coating on the aluminum foil. Sulfuric acid will not work.

Purpose: to show that both strong acid and strong base are very reactive.

Materials: 3.0 mol/L hydrochloric acid, distilled water, 3.0 mol/L sodium hydroxide solution, 3 250 mL beakers, aluminum foil

Method: Place the three beakers on the demonstration bench. Pour 100 mL of the HCl (aq) into one beaker, 100 mL of distilled water into the second and 100 mL of 3.0 mol/L NaOH (aq) into the third.
Make 3 equal size spheres of aluminum foil [about 6” x6” pieces rolled up] and place one into each beaker.

Observe and record what you see.

Observations: *The base will start bubbling first but the acid will catch up. Both bubble and steam vigorously. [Hence: stand back]. Both will have a dark colour in the end [insoluble impurities in the Al foil]. Both will be hot. Let cool when finished. Use one to neutralize the other by pouring one then the other into a 1.0 L beaker with 500 mL of water in it. Dispose of as salt solutions are to be disposed of in your school board. The water does not react.*

Conclusions: Which is safer, acid or base?
Both are dangerous and must be treated with respect when using. Middle [neutral or close to neutral pH ranges are safest. FOLLOW SAFETY INSTRUCTIONS!
What gas is produced?
Hydrogen gas is produced.
Write balanced chemical equations for the reactions that took place.

$6\ HCl\ +\ 2\ Al\ \rightarrow\ 3\ H_2\ +\ 2\ AlCl_3$

$2\ NaOH\ +\ 2\ Al\ +\ 6\ H_2O\ \rightarrow\ 3\ H_2\ +\ 2\ NaAl(OH)_4$ *[sodium aluminate]*

The reactions work because aluminum is an amphoteric metal, in that it can react with both acids [base like] and bases [acid like] due to its location on the periodic table where metals [usually basic] meet non-metals [usually acidic].

Organic Chemistry: An Improved Bromine Test for Saturation.

L. H. Holmes has described how to make chlorine water by mixing 20 mL of concentrated hydrochloric acid with 180 mL of hypochlorite bleach [Javex or Chlorox].[1]If 29 grams of potassium bromide is added to the hypochlorite bleach before the hydrochloric acid is added, a reaction will take place to produce bromine and thus bromine water. This bromine water can be used to test for saturation in carbon compounds. You may cut the amount of acid in half to decrease the concentration of the bromine water. I actually dilute the bleach with an equal volume of water, and use only 3 mol/L HCl and 14 g of potassium bromide to reduce the bromine concentration.

Gloves and goggles are recommended with this preparation.

$HCl\ (aq)\ +\ NaOCl\ (aq)\ \rightarrow\ Cl_2\ (aq)\ +\ NaOH\ (aq)$

$2\ KBr\ (aq)\ +\ Cl_2\ (aq)\ \rightarrow\ Br_2\ (aq)\ +\ 2\ KCl\ (aq)$

The material formed is bromine water, a solution of aqueous bromine and not elemental bromine. Many school boards have banned bromine, but check to find out if bromine water is banned. The two are different. If you can use bromine water, then you can demonstrate the following lab, or if you have adequate ventilation have students perform the following lab.

1. L. H. Holmes,Jr., *J. CHEM. ED.,* **74** , 1326 (1997)

SCH4U1: Lab; Hydrocarbon Reactions, Addition and Substitution

Most hydrocarbon testing labs use hydrocarbons such as cyclohexane and cyclohexene which are volatile and have unpleasant odours. These properties make doing organic chemistry labs in high school unpopular. By substituting water soluble, odourless alcohols propane-1,2-diol, butene-1,4-diol, and butyne-1,4-diol for the hydrocarbons, 1.0 mol/L aqueous solutions of the alcohols can be used to replace the smelly hydrocarbons. The reactions with bromine water are the same and one gets to view the reactions with a triple bond as well as the double and single bonds in organic compounds. A slurry of an aspirin in water is used in place of toluene, for an aromatic hydrocarbon.

Purpose: to observe the rates of substitution reactions in saturated and aromatic hydrocarbons and addition in unsaturated hydrocarbons.

Materials: 4 test tubes, 10 mL graduated cylinder, alkane solution, alkene solution, alkyne solution, aromatic solution, bromine water

Procedure for Testing Hydrocarbons:

1. Take the alkane solution and place 5 mL into a test tube. Add 10 drops of bromine water to the alkane. Note how long it takes the bromine colour to disappear.

What kind of bonds are found in alkanes? [single[
Is the reaction slow [substitution] or fast [addition]? How do you know? [slow, so substitution, bromine colour persists]
Is an alkane saturated or unsaturated? Explain. [Saturated, all carbon carbon single bonds, carbon can not form any more bonds.]

2. Repeat the reaction in part 1 with the alkene solution in place of the alkane solution.

What kind of bond is found in an alkene? [Carbon carbon double bond]
Is the reaction slow or fast? How do you know? [Fast. Bromine colour disappears on contact.]
What kind of reaction took place? [addition]
Is an alkene saturated or unsaturated? Explain. [Unsaturated, the carbon atoms in the double bond can make a single bond and add in other atoms.]

3. Repeat 1 with the alkyne solution in place of the alkane solution.

What is kind of bond is found in an alkyne? [Carbon carbon triple bond]
Is the reaction slow or fast? How do you know? [Fast, the bromine colour disappears quickly.]
What kind of reaction took place? [addition]

Is an alkyne saturated or unsaturated? Explain. [Unsaturated since the carbon atoms in the triple bond can add other atoms to form single bonds.]

4. Repeat the reaction from part 1with an aromatic slurry of crushed aspirin in 5 mL of water in place of the alkane solution.

What is found in an aromatic hydrocarbon? [Benzene or phenyl ring of 6 carbon atoms with supposed alternating single and double bonds]
Is the reaction slow or fast? How do you know? [slow, bromine colour persists]
What kind of reaction took place? [addition]
Is an aromatic hydrocarbon saturated or unsaturated? Explain. [unsaturated since the carbon atoms should all be able to bond to one more atom.]

Also test butter, coconut oil, olive oil and other home materials. [How would you do this?]
For a vegetable oil continue adding bromine water and mixing the two. What happens to the state of the oil as more bromine is added? Explain. *[the oil becomes saturated and solid like a fat would be.]*

Assume that the alkane used is ethane, the alkene used is ethene, the alkyne used is ethyne, and the aromatic compound used is benzene.
Write both word and balanced chemical equations for the reactions that would have occurred.

$CH_3CH_3 + Br_2 \rightarrow CH_3CH_2Br + HBr$

$CH_2CH_2 + Br_2 \rightarrow CH_2BrCH_2Br$

$CHCH + 2\,Br_2 \rightarrow CHBr_2CHBr_2$

$C_6H_6 + Br_2 \rightarrow C_6H_5Br + HBr$

Define: addition reaction for hydrocarbons
Substitution reaction for hydrocarbons
Saturated hydrocarbon
Unsaturated hydrocarbon
Cis-and trans isomers
Aromatic hydrocarbon

The actual compounds used for the hydrocarbons were soluble diols, 1,4-butanediol, 2-butene-1,4-diol, 2-butyne-1,4-diol and 2-acetyl benzoic acid [acetylsalycilic acid]. What are the structures of these compounds? Why were these compounds used in place of the hydrocarbons? [used since less smell and can be dissolved in water.]

At the end of the lab, to get rid of the bromine water add some alkene solution to the bromine water to form the dibromo compound for disposal.

In doing these reactions you will notice that the alkyne reacts more slowly than the alkene. Remember triple bonds are stronger than double bonds so need more activation energy to react. Think nitrogen gas in our atmosphere.

Easy Soap

This activity uses a solid base and a vegetable oil to produce soap easily and relatively safely.

SAFETY NOTE*: Wear eye and hand protection. Solid sodium hydroxide is very caustic.*

Background: Today most people use soap, or another surfactant, for cleaning their clothes, dishes, cars, skin and hair. The first surfactant was soap. Soap making is one of the oldest man made chemical processes. The way people made soap in their homes was to take rendered (melted off meat) fat from animals and placing this in a dish. Wood ashes were then placed on top of the solidified fat and the dish stored for a long time. The bases in the wood ash would react with the fat to produce soap. The ash was removed and soap the soap was cut into bars for use.
The soap made in this way can be tested, students do obtain product. Using a boiling solution of sodium hydroxide with oil does not work as well and produces little product.

The type of chemical reaction for the formation of soap is referred to as a saponification. In this reaction a triglyceride [fat or oil] is reacted with a base such as sodium hydroxide [NaOH] producing glycerol or 1,2,3-propanetriol [$CH_2OHCHOHCH_2OH$] and soap, a salt of a large carboxylic acid such as sodium stearate or sodium octadecanoate [$CH_3(CH_2)_{16}COO^-Na^+$]. The fat or oil is a triglyceride or triester of glycerol. An ester is a compound formed from an alcohol [organic molecule with a hydroxyl or OH group] and a carboxylic acid [an organic molecule with a carboxyl or COOH group] linking together with the formation of water. Since glycerol has three hydroxyl groups, the molecule can react with three carboxylic acid molecules to form the triglyceride. In this experiment, it will be assumed that the vegetable oil used [olive oil works well] is glycerol trioleate or 1,2,3-propanetriyltrioctadecenoate [$CH_2\{CH_3(CH_2)_6CH{=}CH(CH_2)_6COO\}CH\{CH_3(CH_2)_6CH{=}CH(CH_2)_6COO\}CH_2\{CH_3(CH_2)_6CH{=}CH(CH_2)_6COO\}$].

Three moles of sodium hydroxide are required to react with one mole of the triglyceride, three moles of the soap, sodium trioleate and one mole of glycerol are produced:

3 NaOH +
$CH_2\{CH_3(CH_2)_6CH{=}CH(CH_2)_6COO\}CH\{CH_3(CH_2)_6CH{=}CH(CH_2)_6COO\}CH_2\{CH_3(CH_2)_6CH{=}CH(CH_2)_6COO\}$
$\rightarrow$ 3 $CH_3(CH_2)_6CH{=}CH(CH_2)_6COO^-Na^+ + CH_2OHCHOHCH_2OH$.

In a manner similar to the pioneers, you will use a triglyceride but this time an oil, and in place of the wood ash, an active ingredient of the wood ash, granular sodium hydroxide. The granular sodium hydroxide is more concentrated than the bases in wood ash, so that the reaction will not take months, but only hours.

Activity:

Place 10.0 mL of olive oil into a small plastic cup.
Add 1.22 g of granular sodium hydroxide.

Let this mixture stand until the next class.
Note any changes that have taken place in the material in the cup.

Testing the Product:

Add a small amount [what will fit on the tip of a wood splint] of the soap to 5 mL of distilled water in a test tube. Stopper the test tube and shake the contents. What do you observe? Add red and blue litmus strips. Is the solution acidic, neutral or basic?

Repeat with tap water.

Repeat with 0.01 mol/L calcium chloride solution [hard water].

Add 5 drops of vegetable oil to 5 ml of distilled water in a test tube. Shake the contents. How long does it take the oil to separate out of the water?

Add 5 drops of oil to 5 mL of distilled water in a test tube, and also add a small amount of the soap. Stopper the test tube, shake the contents Time how long it takes for the oil to separate.

Questions:

1. Verify the quantities of reactants used in this experiment by calculating the moles glycerol trioleate [olive oil] used, the moles of sodium hydroxide required to react with the olive oil, and the mass of this number of moles of sodium hydroxide. The density of glycerol trioleate is 0.8988 g/ mL.

2. How have the reactants changed?

3. What does the litmus test indicate about the water soap solution?

4. What effect does shaking the mixture have?

5. What happened to the mixture in the calcium chloride solution? Explain.

6. Is the water in your school hard or soft? What test allows you to determine this?

7. What are immiscible liquids? Give an example from this lab.

8. What are miscible liquids? Give an example. Try mixing these.

9. What are emulsifying agents? Give an example from this lab.

10. Explain the observations with the oil and water, and the oil, water and soap.

11. Explain how emulsifying agents work.

12. Why is “Draino” [sodium hydroxide is its main ingredient] used to unclog drains?

Teacher’s Notes

This exercise can be used in the organic chemistry section of your course, to review stoichiometry, look at hard water, reaction rates, consumer chemistry, set up experiments to test for the best product i. e. to do some scientific enquiry.
This experiment provides a better method of producing soap than the heating of fats or oils with concentrated sodium hydroxide solution, sometimes with ethanol, and trying to salt out some soap that is found in most soap making labs. The students spend a period trying to make soap with poor results with the old method. The method described here guarantees a product, with about 10 min spent at the end of one period, the students can spend a lab period testing their product.
Any liquid vegetable oil can be substituted for the olive oil. Cheap olive oil works just as well as expensive olive oil. You can have the students use various oils and determine which give the “best” soap. The criteria can be foam height, how long the foam lasts, how long the soaps keep oil emulsified, appearance, texture, odour, or some combination of these. If you wish to use a solid triglyceride [animal fat or solid vegetable oil such as coconut oil] the fat should first be melted in a hot water bath at ~60^{o} C. [see below]
The sodium hydroxide must be granular, flakes or pellets do not react fast enough. Buy the smallest quantity available since once open the sodium hydroxide will react with water and carbon dioxide from the air to form sodium carbonate and sodium hydrogen carbonate which will not saponify the oils next year. Do not grind sodium hydroxide pellets or flakes in a mortar with a pestle since the solid is too caustic to make this a safe procedure.
For mixing the ingredients for soap use 30 mL graduated mixing cups available from hobby stores. These are disposable. If you use a test tube or beaker the soap is very thick and difficult to clean out. I dispose of the left over material by placing the soap and cups in a plastic bag and placing this in the regular garbage.

You can study factors that affect the rates of reaction with this experiment. To look at surface area, follow the instructions on the student page, but also set up a cup with the correct mass of sodium hydroxide pellets. The pellets will not form soap over night. Concentration can be studied by using wood ashes on the oil and seeing how long it takes for the soap to form. The less concentrated base in the ashes takes more time. Temperature effects can be studied by setting up water baths of 40^{o} C and 60^{o}C and comparing the rates at these temperatures with the rate at room temperature. By adding a drop or two of water and stirring, the effect of a catalyst can be simulated.

To make the hard water, 0.01 mol/L $CaCl_2$ (aq.) solution dissolve 1.47 g of calcium chloride dihydrate in 1.00 L of water.

Answers to Questions:

1. 10 mL of olive oil translate into 10 mL x 0.8988 g/ mL = 8.988 g of glycerol trioleate.

At 885.5 g/mol this is 8.988g/885.5 g/mol = 0.01015 mol.
This requires 3 mol NaOH for each mole of oil so 3 x 0.01015 mol = 0.03045 mol NaOH is needed.
0.0315 mol NaOH is [40.0 g/mol] 40.0 g/mol x .03045 mol = 1.218 g ~1.22 g [what can be weighed out on most high school balances]

2. The oil and white solid will form a thick slippery gel.

3. Red litmus will turn blue indicating a basic solution.

4. A foam forms on the liquid as one would expect for a soap.

5. Calcium ions form an insoluble soap, calcium oleate which shows up as a white precipitate or curd. Calcium ions, magnesium ions, iron ions or acidic hydronium ions will cause the same effect. If these ions are present in the water the water is termed hard. One effect of hard water is its precipitation of soap rendering the soap useless for cleaning. [You may wish students to do some research about hard and soft water.]

6. If the tap water forms a white precipitate to any extent, and less foam than the distilled water, this is an indication that the water has some degree of hardness. If the tap water is as foamy as the distilled water, then the water in the school is soft.

7. Immiscible liquids are liquids that do not dissolve in one another such as oil and water.

8. Miscible liquids are liquids which do dissolve in each other in all proportions. Ethanol and water are examples.

9. An emulsifying agent is a material that can hold two immiscible liquids together so that they do not separate. Soap is an example of an emulsifying agent that can hold oil and water together. Egg yolk is used to hold vinegar and oil together

in mayonnaise. [Students may wish to find a recipe for mayonnaise and try to make some as an addition to this lab.]

10. The oil does not mix with the water because the two are immiscible. The oil remains in the water, [and may take a very long time to separate, so ask students to stop timing after about 1 or 2 minutes with the soap, oil water mixture] with the soap since soap emulsifies or holds the two liquids together.

11. Liquids are immiscible because they are of different polarities. Oil is non polar while water is polar. Emulsifying agents such as a soap are both polar and non polar. The polar part sticks to water while the non polar part sticks to the oil. The oil soap ends will stick together in small bundles or micelles with the polar ends sticking to the water. The micelles are thus suspended in water and the two liquids are "mixed".

12. The "Draino" supplies sodium hydroxide which reacts with oils and fats in the clog making them more water soluble and able to move through the drain pipe.

1. A version of this was published in CHEM 13 NEWS October 2003.

2. For a detailed explanation of how soaps work see David Katz's web site www.chymist.com

3. Chem 13 News.

Easy Reverse Soap

A normal soap is a salt of a large carboxylic acid, such as sodium oleate (cis-9-octadecenoate) $CH_3(CH_2)_7CH{=}CH(CH_2)_7COO^-Na^+$. The long hydrophobic end that will stick to oil is the hydrocarbon end, while the charged ionic end will be attracted by water allowing the soap to emulsify oil and water.

What if instead of a large hydrophobic negative countered by a small positive ion, a large hydrophobic positive ion with a small negative ion were produced? What would this material behave as? This material would be a ***reverse soap*****,** i.e. the ion charges are reversed from a normal soap.

You can demonstrate the idea by taking a glass stopper with a drop of concentrated ammonium hydroxide on it next to a glass stopper with concentrated hydrochloric acid on it. A white smoke or precipitate of ammonium chloride ($NH_4^+Cl^-$) will appear.

A group of organic amines R_3N will undergo similar reactions with hydrochloric acid. In Ontario the grade 12 college chemistry preparation course has a section to study properties of amines. This lab can be used for the study of a chemical property of amines. They do form salts with mineral acids. If N,N-dimethyldodecyl amine is used,

$(CH_3)_2CH_3(CH_2)_{11}N$, [available from Aldrich chemical company, a reverse soap can be produced.

The dimethyldodecylamine is a liquid at room temperature. It does exhibit the characteristic foul odour of an amine, so work in a fume cupboard. Concentrated hydrochloric acid also has a very strong odour. Place 10 mL of the amine in a glass jar. Add with mixing, 2.7 mL of concentrated hydrochloric acid. A white smoke similar to that seen with the ammonia and hydrochloric acid will be observed in the jar. The mixture becomes noticeably warm, an exothermic reaction. The amount of dimethyldodecylamonium chloride, $(CH_3)_2CH_3(CH_2)_{11}NH^+Cl^-$, formed should be enough for a class of 20 to complete the tests with the soaps. A clear gel forms.

For each test take the amount of material that will fit on the end of a wooden splint and dissolve this in 5 mL of distilled or deionized water in a 20 mL test tube. Stopper the test tube and shake the contents to dissolve the soap.

Compare the sodium oleate to the dimethyldodecylammonium chloride with the following tests.

Do the soap solutions foam on shaking?

Add a few drops of bromthymol blue indicator to each solution. The regular soap will be basic, while the reverse soap is acidic. (You can also measure the pH of each solution.)

The carboxylate ion hydrolyzes according to:

$$RCOO^- + H_2O \rightarrow RCOOH + OH^- \text{, a basic solution,}$$

While the ammonium ion hydrolyzes as:

$$R_3NH^+ + H_2O \rightarrow R_3N + H_3O^+ \text{, an acidic solution.}$$

Compare the ability to emulsify oils by adding five drops of a vegetable oil to each test tube and agitate. Do the soaps keep the oil suspended in the water, an emulsion, or does the oil separate out and float to the top? A good soap should form an emulsion.

Add a sample of each soap to a separate sample of hard water, 5 ml of clear limewater will do. Stopper and agitate the mixture. A regular soap will form an insoluble calcium soap, removing its ability to emulsify oil. Add 5 drops of a vegetable oil and try each hard water mixture for emulsifying ability. Which soap, normal or reverse, works best in hard water?

Now mix up one more 5 mL sample of each soap in 5 mL of distilled water. Mix the two solutions. Again a precipitate may form. This time the two large ions will form a large less soluble salt, dimethyldodecylamonium oleate, $(CH_3)_2CH_3(CH_2)_{11}NH^+$ $CH_3(CH_2)_7CH{=}CH(CH_2)_7COO^-$, and sodium chloride.

What would a reverse soap be used for? You probably have some in your home. Some fabric softeners have a material called dimethyltallow ammonium chloride, which is the material you have just made. The reverse charges on the fabric softeners counter the charges that build up on clothes with the normal soaps during washing. Similar materials are also used in hair conditioners. The reverse soaps also have anti-bacterial properties as do regular soaps.

Instant Soap

A normal soap is a salt of a large carboxylic acid, such as sodium oleate (cis-9-octadecenoate) $CH_3(CH_2)_7CH{=}CH(CH_2)_7COO^-Na^+$. The long hydrophobic end that will stick to oil is the hydrocarbon end, while the charged ionic end will be attracted by water allowing the soap to emulsify oil and water.

Sodium oleate can be made easily by mixing 10 mL of oleic acid (0.0315 mol) with 10.5 mL of 3.0 mol/L NaOH (aq.) or 3.0 mol/L KOH (aq) in a disposable container. You should have oleic acid in the school. A lab in the old physics curriculum used oleic acid to produce a thin film on water. The thickness of the film, and the order of magnitude of molecular size was calculated using the volume of the oleic acid drop divided by the area of the film. There may be some hidden away in the physics lab if you have none in the chemistry lab. You can also order oleic acid from any science supplier.

Use distilled water in making up the base solutions and distilled water to show that the soap does form suds.

Bubble Molecules

Try this one to obtain the answers.

The Bubble Solution:

Mix 80 mL of distilled water with 15 mL of Joy™ or Dawn™ dish detergent and 5 mL of pure glycerine. Allow the solution to mix thoroughly. The solution improves with age. You can experiment with this mixture by varying the amounts of ingredients. [Thanks to David Katz]

The VSEPR Model with Bubbles:

Purpose: To simulate the shapes of molecules as predicted by the VSEPR theory with soap bubbles.

Materials: Soap bubble solution, Beaker, Straw, Protractor, Flat hard surface, Paper towels.

CAUTION: **Wear goggles. Soap bubbles can cause eye irritation.**

Procedure: 1. Pour 25 mL of soap bubble solution into a 100 mL beaker.
2. Wet a 10 cm by 10 cm square of the hard surface with soap bubble solution.
3. Dip the straw into the bubble solution and blow a bubble on the wet surface. Blow a second bubble on the surface adjacent to the first bubble so that they touch. The bubbles should be the same size.

The bubbles the electron clouds attached to a central atom in the molecule. What shape of molecule results when two bonding pairs only are around the central atom? Give an example of such a molecule. What is the angle between the centers of the atoms?

3. Repeat but with three touching atoms of the same size. Measure the angle formed where the bubbles meet.

What shape will three atoms around a central atom with three bonding pairs of electrons give? Give an example of a molecule with this shape. What other shape is associated with this electron pair arrangement? Give an example of a molecule with this shape. What happens to the angle with this shape?

4. Now enlarge one of the three bubbles then decrease the size of this bubble.

What happens to the angle between the atoms? What happens to the angle between atoms in molecules as there is a switch from bonding to non- bonding electron pairs?

5. On top of the three bubbles of the same size, blow a fourth bubble where the three bubbles meet. Measure the angle that is formed between the bubbles.

What shape is formed in this case? Give an example of a molecule with this shape. How many bonding pairs are associated with the central atom? What angle forms between atoms in this shape? What other molecular shapes are associated with this electron pair arrangement? Give examples of molecules with these shapes.

Parts 6 and 7 require much coordination and patience between students.

6. Have five partners blow five bubbles of the same size at the same time to join and to float in the air. This takes practice.

What shape do these bubbles arrange themselves into? Give an example of a molecule with this shape. What other shapes are associated with this electron pair arrangement? Give examples of molecules with these shapes. What angles are associated with these molecules?

7. Have six partners blow six bubbles of the same size into the air at the same time to join and float in the air.

What shape do these bubbles arrange themselves into? Give an example of a molecule with this shape. What other shapes are associated with this electron pair arrangement? Give examples of molecules with these shapes. What angles are associated with these molecules?

8. To clean up first wipe the soap bubble areas with **DRY** *paper towel. Then wipe the area using vinegar on a paper towel. Dry the area.*

Conclusion: *Write the names, formula, Lewis electron dot diagrams, and shapes of all the molecules named above. You can arrange these in a table format. Associate the molecules with the same number of electron pairs around the central atom but different shapes, into the same sections.*

Demo or Lab; The Reaction Between Calcium and Water

This lab was adopted from my high school lab manual used in the 1960's in Ontario, "Experiments in Laboratory Chemistry" by Croal, Couke and Louden, Copp Clark Publishers.

Purpose: to investigate the reaction between calcium and water, and to identify the products and some of their properties.

Materials: a large beaker, water, calcium turnings, file, test tubes, wood splint, match, funnel, filter paper, straw, litmus paper, sand paper, magnesium ribbon

Procedure:

1. Take a piece of calcium turning. File off the outer coating. Do not touch the calcium underneath! *Describe the coating. Describe the calcium. What is the coating.*
2. Place the large beaker in the sink. Fill it with water. Fill with water and invert full of water, three test tubes into the beaker.
3. Place the calcium into the water and collect two test tubes full of the gas that is produced. Place these on the lab bench mouth down, keeping them mouth down at all times.
4. The third test tube allow to fill only 1/8 to ¼ full before removing and allowing air to replace the water. Keep this test tube separate, but also mouth down on the bench.
5. Take the two full of product gas test tubes and keep one mouth down while the other is mouth up. Light a wooden splint. Insert the lit splint into the inverted test tube. Observe very closely what happens. Then withdraw the wooden splint and again observe what happens. Now insert a burning splint into the mouth up test tube. Observe what happens this time. *What do the results tell you about the flammability of the gas? What do the results tell you about the ability of the gas to support combustion? What do the results tell you about the density of the gas?*
6. Keep the 1/8 full of product gas mouth down and insert a burning wooden splint. *Describe the result. What was the colour of the flame. Explain the difference*

between the results with the full and 1/8 full test tube. What gas is produced in the reaction between calcium and water?

7. Fold a filter paper, place it in a funnel, and place it on a test tube. Pour about 15 mL of the liquid in the beaker into the funnel and collect about 3 cm deep of filtrate in the test tube.
8. Test the filtrate with litmus paper. *What does the test tell you?*
9. Take the straw and gently blow your breath into the filtrate until change occurs. *What gas is added from your breath that is less plentiful in air? What substance is the filtrate? Describe the solubility of the substance in the filtrate.*
10. Write a word equation and a balanced chemical equation for the reaction between calcium and water.
11. Write a word equation and a balanced chemical equation for the combustion of the gas produced.
12. Write a word equation and a balanced chemical equation for the reaction between the gas in your breath and the material in the filtrate.
13. The following will be set up as a demonstration to go over a week. A piece of magnesium ribbon at least 30 cm long will be cleaned with sand paper. *Describe the clean magnesium. What did the sand paper remove from the magnesium?*
14. The magnesium will be rolled into a spiral and placed at the bottom of a beaker of water. A funnel will be placed over the magnesium ribbon, and a test tube full of water inverted into the water in the beaker over the funnel stem.
15. After at least a week, observe the magnesium and the test tube. *What changes have taken place?*
16. When enough gas has collected in the test tube, keeping the test tube inverted, remove the test tube and insert a burning wooden splint. *Describe what happens.*
17. Write the word and chemical equation for the reaction between magnesium and water.
18. ***Your teacher may demonstrate the effect of steam on a burning splint and a burning piece of magnesium.*** *What implications does this have regarding fire fighting?*
19. Compare the reactivity of magnesium and calcium. Give a reason for the difference in reactivity.
20. Predict the reactivity of strontium and barium with water. How would the reactions compare to that of calcium and water? Explain your answer.
21. How would barium be stored? Explain.
22. Compare the reactivity of magnesium with sodium and calcium with potassium. Explain the differences.

The key reaction here is the production of hydrogen gas from the reaction between calcium metal and water:

$$Ca + 2\,H_2O \rightarrow H_2 + Ca(OH)_2$$

The Reaction Between Copper(II) Ions and Aluminium

The late Cliff Schroder liked this reaction because it shows a single displacement reaction or redox reaction that is catalyzed by the presence of chloride ions, or bromide ions.

Pat Funk reworked this into a lab to show the mole relationship in the reaction.

RELATING MOLES TO COEFFICIENTS OF A CHEMICAL EQUATION

Purpose: to determine the mole ratio of aluminum used to copper formed in a single replacement reaction.

Materials: make a list of all materials used.

Procedure: 1. Obtain a piece of aluminum foil that has a mass of 0.54 gram. Cut it into small pieces and place it into a clean dry 250 mL beaker. Predict the mass of copper that should form from the 0.54 g aluminium.
2. Dissolve 10 g of copper(II) sulphate pentahydrate in 100 mL of distilled water. Add this solution to the 0.54 g of aluminium.
3. Add a pinch of sodium chloride as a catalyst.
4. Place the beaker on heater and begin to heat it. Use a stirring rod to regularily stir the mixture until all the aluminum foil has dissolved. Do not allow the mixture to go above $60^{o}C$. Remove the heat source if necessary.
5. Allow the mixture to sit for at least five minutes.
6. Obtain the mass of a filter paper.
7. Filter the copper onto the filter paper.
8. Wash the solid with 10 mL of distilled water five times.
9. Do a final wash of the solid with 10 mL of acetone.
10. Place the filter paper on a watch glass to dry.
11. When the copper has dried obtain the mass of the filter paper plus copper.

Observations: 1. Describe the changes that take place in the beaker.
2. Describe the precipitate and supernatant.
3. What masses were measured? What are the masses?

Calculations: 1. Calculate the number of moles of aluminum used.
2. Calculate the mass and number of moles of copper formed.
3. Calculate the mole ratio of copper/aluminium.

Questions: 1. Write a balanced chemical equation for the chemical reaction which took place.
2. Compare your experimental result with the theoretical result. Explain why they do or do not agree.
3. What colour was the supernatant at the end of the experiment? Explain.

Conclusion: What did you learn concerning mole ratios, balanced chemical equations, and the atomic theory from doing this experiment?

New vocabulary: Define and /or explain, with examples where appropriate , the following:

1. coefficient
2. aqueous
3. single replacement reaction
4. limiting reagent
5. excess reagent
6. catalyst

P. E. FUNK, Retired, Pataskala, OH

The idea is also used to set up a pop can to rip in half. Most lab outlines have you use copper(II) chloride solution for the can rip, however copper(II) sulfate with sodium chloride added will work just as well and costs less. Save the solution since it can be used several times before not working.
Cliff Schroder would ask his students to think of another catalyst. Most would use other chlorides, but remember bromides should also be able to disrupt the aluminium oxide coating on the pop can. Try this to see if it works.

The Great Pop Can Rip

1. Score the inside of a pop can with a metal file.
2. Fill the can with copper sulfate solution to just above the score mark.
3. Add a pinch of sodium chloride (salt) to the copper sulfate solution. The teacher will set up a can without any salt added.
4. Place the can inside a beaker.
5. When the can just begins to leak, pour the contents out of the can and rip the can in two. ***CAUTION! The edges of the torn can are sharp. Wear protective gloves.***
6. Clean the beaker and put the can in the garbage.

Explain what happened to the aluminum can.

What was the purpose of adding the salt [sodium chloride] to the copper sulfate solution?

What else besides salt could be added to the copper sulfate solution to give the same result?

Demo: Reversible Redox

This was adapted from “The Joy of Chemistry” byCathy Cobb and Monty L. Fetterolf, Prometeus Books.

If a chloride and a bromide work to remove the aluminum oxide coating, will an iodide work? Not in the presences of copper(II) ions. The oxidation reduction potentials are such that the iodide ion will give electrons to the copper(II) ions to form copper and iodine. The copper reduction goes through a two step process to copper(I) then copper.

$Cu^{+2} + 2\ e^- \rightarrow Cu$
$2\ I^- \rightarrow I_2 + 2\ e^-$

If 0.2 mol/L potassium iodide solution is mixed with an equal volume of 0.2 mol/L copper(II) sulfate solution in a test tube, the two mix to form a red brown opaque slurry. Add mineral oil to the mixture. Stopper the test tube and shake the mixture and you will see the presence of elemental iodine in the oil as the colourless oil turns violet. [A small sample can be taken and rubbed on paper to show a dark violet blue colour, indicative of iodine on starch. [What is paper coated with?]

In the Joy of Chemistry it is suggested that ammonia solution be added. Save the bad odour for elsewhere and instead add potassium sodium tartrate solid or sodium citrate solid and agitate the mixture. The purple in the mineral oil begins to disappear and the solution begins to turn blue, the colour of copper(II) ions. The system is reversing. According to Le Chateliere's principle, the added tartrate will complex, thus remove the few copper(II) ions remaining. The system will reverse direction in order to produce more copper(II) ions and thus reduce the concentration of iodine.

Demo: More on Copper(II) Ion Complexation

Add 0.2 mol/L Na_2CO_3 (aq) to an equal volume of 0.2 mol/L $CuSO_4$ (aq) in a test tube and the solution becomes cloudy as a precipitate of $CUCO_3$ (s) and $Cu(OH)_2$ forms.

Now repeat this but before mixing the two add 0.4 mol/L sodium citrate solution to an equal volume of copper(II) sulfate solution. Note the change in intensity of the blue colour as the copper(II) citrate complex forms.
Now add the sodium carbonate solution. No precipitate forms since the copper(II) ions are surrounded by citrate ions.
This idea is used for testing reducing sugars such as glucose. The reducing sugars reduce copper(II) to copper (I) ions but this must be done in a basic medium. To keep the copper(II) ions from precipitating out, they are complexed with the citrate ion [Benedict's solution] or tartrate ions [Fehling's solution].

Lets add some glucose to some of the basic copper(II) citrate solution and heat it in a hot water bath to see if we do have Benedict's solution. The change in colour to a red brick indicates a reducing sugar with copper(II).

Design a Buffer: Background

I worked with John Dragert at Stouffville DSS in designing this exercise. John is still teaching at the school.

Questions to gain student interest:

How does blood maintain a pH of 7.2 – 7.4?
How are swimming pools kept at a pH of 7.2 – 7.6?

Problems for students to consider:

Add 0.500 mL of 0.100 M HCl to 100.0 mL of distilled water at pH 7.00. What is the final pH?

Moles acid / volume = (0.5 mL x 0.10 mol/1000 mL) / 0.1005 L = 0.0005 M
pH = - log [H^+] = - log 0.0005 = 3.30
A small amount of acid causes a large pH change.
Add 0.500 mL of 0.100 M NaOH to 100.0 mL of distilled water at pH 7.00. What is the final pH?

pH = 14 – pOH = 14 – 3.30 = 10.7.

Again a small amount of base creates a large shift in pH.
Have the class try this.

What to consider:

Enter weak acid weak base equilibrium for an explanation.
Example: acetic acid acetate ion equilibrium, K_a = 1.80 x 10^{-5}.
Simplified equilibrium equation:

$$CH_3COOH = CH_3COO^- + H^+$$

$K_a = [H^+] \cdot [CH_3COO^-] / [CH_3COOH]$

Can a solution with pH 7.00 be made with this acid base system?

Substitute into K_a expression:
$1.80 \times 10^{-5} = (10^{-7}) \cdot [CH_3COO^-] / [CH_3COOH]$
Assume [CH_3COOH] = 0.100 M
Solve for [CH_3COO^-] = 18 M.
A very large difference in conjugate acid – base concentrations.
What pH makes better sense to attempt to make a solution for? Why?

Try pH of 5, then ratio of [base] / [acid] = 1.80.
If [acid] = 0.100 M then [base] = 0.180 M.
Mix 0.0180 mole x 82 g / mole = 1.48 g $NaCH_3COO$ with 100 mL 0.100 M CH_3COOH, to make a pH 5 solution.

Set up class to make the solution.

Add 0.500 mL 0.100 M HCl, check pH, add 0.500 mL 0.100 M NaOH, check pH What happens? Why? Both acid and base present to react with added acid and base maintaining pH.
[acid] on adding acid changes to 0.105 M, [base] changes to 0.175 M,
[H^+] changes to 1.80 x 10^{-5} x .105 / .175 =1.08 x 10^{-5}

$pH = -\log 1.08 \times 10^{-5} = 4.97$
Similarly on adding base the pH becomes 5.03.

Now have students design their own buffer:

Possible Buffer Systems:

$KHSO_4$	K_2SO_4	$K_a = 1.20 \times 10^{-2}$
$NaHCO_3$	Na_2CO_3	$K_a = 5.61 \times 10^{-11}$
$NaHSO_3$	Na_2SO_3	$K_a = 1.02 \times 10^{-7}$
KHC_2O_4	$K_2C_2O_4$	$K_a = 6.40 \times 10^{-5}$
K_2HPO_4	K_3PO_4	$K_a = 2.2 \times 10^{-13}$
KH_2PO_4	K_2HPO_4	$K_a = 6.23 \times 10^{-8}$

Take simple calculation, not complex calculation involving coupled equilibria.

What systems are used to control pH as buffers?

K_a HclO = 2.95×10^{-8}, used for swimming pool pH. {comments?}

K_a HbrO = 2.06×10^{-9}, used for hot tub pH. {comments?}

Ka $H_2CO_3 = 4.30 \times 10^{-7}$, blood buffer system, good for pH 7

Student Lab: Design a Buffer

Introduction:

Buffers are chemical systems that resist changes in pH due to the addition of acids or bases. They consist of an acid and its conjugate base. For instance one buffer system could be:

Acetic acid and Sodium acetate

The conjugate base is added as a salt. In this case, a sodium salt (Sodium acetate) is used. The buffer has an optimal pH range centred on the negative log of the Ka value. In today's lab, you will be designing a buffer system and testing it's properties.

Hypothesis:

State how your buffer should work in terms of:

a) resistance to pH change
b) pH range it will work the best in

Materials:

Write your own complete materials list. For equipment include the number of each piece of equipment use (i.e., 1 retort stand). For glassware include the number of each piece
used and the size used (i.e., 2 400 ml beakers).

Procedure:

(In Report, refer to handout but explain any changes made)

1. Mix a 0.1 mol/L solution of your acid in a 100 ml volumetric flask
2. Dissolve 0.18 mol of your given salt into the acid in order to make a buffer
3. Calculate the pH of the buffer at equilibrium
4. Measure and record the actual pH of the buffer.
5. Pour 10 ml of the buffer solution into a small beaker.
6. Add 1 ml of 0.1 mol/L HCl. Record the pH. Add 9 ml of 0.1 mol/L HCl, for a total of 10 ml of acid added and record the pH again.
7. Take a second 10 ml portion of the buffer out of the volumetric flask.
   Add 1 ml of 0.1 mol/L NaOH to this portion and record the pH. Add 9 ml of 0.1 mol/L NaOH for a total of 10 ml of base added and record the pH again.
8. Obtain 10 ml of distilled water. Add 1 ml of 0.l mol/L HCl and record the pH.
9. Obtain 10 ml of distilled water. Add 1 ml of 0.1 mol/ L NaOH and record the pH.
10. Dilute the original buffer to 1/10 of the original concentration. Measure 10 ml of the original buffer solution. Clean out the volumetric flask. Pour the 10 ml of buffer back into the flask and add 90 ml of distilled water (i.e., fill to the line) Record the pH.
11. Add 1 ml of 0.1 mol/L HCl to 10 ml of the diluted buffer. Record the pH. Add 9 ml of 0.1 mol/ L HCl, for a total of 10 ml of acid added and record the pH.
12. Add 1 ml of 0.1 mol/L NaOH to 10 ml of the diluted buffer. Record the pH. Add 9 ml of 0.1 mol/ L NaOH, for a total of 10 ml of base added and record the pH.

Possible Buffer Systems (Circle the one you are assigned).

$KHSO_4$	K_2SO_4
$NaHCO_3$	Na_2CO_3
$NaHSO_3$	Na_2SO_3
KHC_2O_4	$K_2C_2O_4$
K_2HPO_4	K_3PO_4

Questions

1. Correctly define the term buffer and explain how a buffer resists changes in pH.
2. Which is the acid and which is the conjugate base in your system? How do you know?
3. Why don't the calculated and measured pH values match for your buffer system?
4. How does diluting the buffer affect the buffer capacity (amount of acid or base the buffer can protect against)? Explain why.
5. Why are swimming pools and hot tubs buffered?

6. Your body needs to keep blood in a pH range of 7.1 to 7.4. If your blood pH falls too far out of that range you will die. How does your body keep blood in the proper pH range?

Permanganate Plus Oxalate

In reaction rate and redox labs one will often find a reaction between potassium permanganate and oxalic acid. The permanganate solution is usually 0.01 mol/L and the oxalic acid concentration is 0.20 mole/L. The net ionic equation for the reaction is:

$$16\ H^+ + 2\ MnO_4^- + 5\ C_2O_4^{-2} \rightarrow 10\ CO_2 + 2\ Mn^{+2} + 8\ H_2O$$

The rate lab uses this reaction as a self catalyzing reaction since the manganese(II) ion is a catalyst for the reaction. Add $MnSO_4$ (aq) and the reaction speeds up. Even though the concentration decreases as more permanganate is added the reaction rate is faster.

However the concentrations are so low that one sees the disappearance of the permanganate purple, but not the formation of the carbon dioxide.

If 3.16 g of potassium permanganate and 6.30 g of oxalic acid dihydrate are mixed in a 250 mL Erlenmeyer flask inside a one liter beaker, cover the solid with distilled water and 30 mL of 3 mol/L sulfuric acid [excess] is added, the evolution of carbon dioxide can be observed. The solid mixture cannot be left standing since it is an organic material in contact with an oxidizer.

Test the gas produced to show it is carbon dioxide. The pale pink solid remaining is manganese(II) sulfate.

Have the students check the stoichiometry of the reaction mixture.

Creaming the Teacher: *[Safety from George Hague]*
Have a school administrator hit you in the face with a whipped cream or shaving cream pie the first week of classes. Have your safety goggles on, a lab coat on, a towel around your neck, do not inhale as the pie strikes your face. Your face and goggles are covered with the cream, but as you remove your goggles your eyes and the area around the eyes are cream free. If there is a chemical accident your eyes will be safe if you wear your goggles.

Use of Paper Clips:

1. ***Chemical Formulas: [Various Presenters]*** Use coloured paper clips. Assign a colour to each main group on the periodic table such as red for oxygen [the chalcogens], black for carbon, white for hydrogen and the alkali metals, green for fluorine [the halogens] etc. You can have the students place the paper clips on a periodic table. The association with the family of elements results in the specific paper clip having a specific valence or charge. For example red oxygen will be -2, green halogen -1, white hydrogen +1, etc. Now have students put paper clip

formulas together such as two white on one red, a white with a green and so on. Note that the $(\text{white})_2\text{red}$ represents any of the formulas that can form from hydrogen or an alkali metal and any member of the oxygen family. After a few trials with these two families of elements [sodium sulfide, hydrogen telluride, potassium oxide, rubidium selenide] the patterns should begin to emerge for the students.

Activity; Formulas of Compounds

Purpose: to understand the formulas of compounds.

Materials: paper clips, periodic table.

Method: 1. Obtain 7 different colours of paper clips.
2. Assign one colour of paper clip to each of the families of main group elements.
3. What is the charge of the ion for each of the main groups of elements?
4. Assign that charge to the paper clip for each group.
5. Combine the paper clips to make a "neutral molecule."

6. Make "paper clip molecules" for the following compounds: sodium chloride, potassium fluoride, lithium bromide, rubidium iodide, cesium astatide.
What do you notice about the formulas for the compounds?
Write the formulas for each compound.

7. Make "paper clip molecules" for the following compounds: sodium oxide, potassium sulfide, lithium oxide, rubidium sulfide, cesium selenide.
What do you notice about the formulas for the compounds?
Write the formulas for each compound.

8. Make "paper clip molecules" for the following compounds: beryllium chloride, magnesium fluoride, calcium bromide, strontium iodide, barium astatide.
What do you notice about the formulas for the compounds?
Write the formulas for each compound.

9. Make "paper clip molecules" for the following compounds: beryllium oxide, magnesium sulfide, calcium selenide, strontium oxide, barium sulfide.
What do you notice about the formulas for the compounds?
Write the formulas for each compound.

10. Make "paper clip molecules" for the following compounds: boron chloride, aluminium fluoride, gallium bromide.
What do you notice about the formulas for the compounds?
Write the formulas for each compound.

11. Make "paper clip molecules" for the following compounds: boron oxide, aluminium sulfide, gallium selenide.
What do you notice about the formulas for the compounds?

Write the formulas for each compound.

12. Make “paper clip molecules” for the following compounds: carbon tetrachloride, silicon tetrafluoride, germanium tetrabromide.
What do you notice about the formulas for the compounds?
Write the formulas for each compound.
13. Make “paper clip molecules” for the following compounds: carbon disulfide, silicon dioxide, germanium diselenide.
What do you notice about the formulas for the compounds?
Write the formulas for each compound.

14. Make “paper clip molecules” for the following compounds: sodium nitride, potassium phosphide, lithium arsenide.
What do you notice about the formulas for the compounds?
Write the formulas for each compound.

15. Make “paper clip molecules” for the following compounds: beryllium nitride, magnesium phosphide, calcium arsenide.
What do you notice about the formulas for the compounds?
Write the formulas for each compound.

16. Make “paper clip molecules” for the following compounds: boron nitride, aluminium phosphide, gallium arsenide.
What do you notice about the formulas for the compounds?
Write the formulas for each compound.

17. Make general statements about writing formulas for binary compounds.

18. Make general statements about writing names for binary compounds.

19. Please return the paper clips to their proper container.

2. ***Limiting and Excess Reagents: [Pat Funk]*** Give students random samples of large and small paper clips. Have them build as many $[\text{small}]_2$ large models as they can with the paper clips they have. Are any left over? Which kind? Which were in excess? Which type of paper clip was limiting? Relate to chemical reactions. Repeat with other combinations.

3. ***Equilibrium: [Pat Funk and Irwin Talesnick]*** Have two students build a model each of a $[\text{small}]_2$ large paper clip “molecule” on a given word say “go” from a sample of the two types of paper clips. At the same time have one student take apart one of the models apart [the first time there are no models to take apart.] Each time the students are instructed to “go” they repeat the process. Eventually there are a certain number of product “molecules’ an excess of one “reagent” and only one molecule is being made and taken apart at one time. That is equilibrium is reached. Adding some more of the limiting reagent will cause more product to

form and shift the equilibrium. Removal of either the limiting or all the excess plus at least one more of the excess “reagents” will result in less product. To increase temperature add one more person to each side so that three “molecules” are made and two are taken apart. The equilibrium and equilibrium rates will shift. Try other combinations of students and paper clip “molecules.”

4. ***Addition Polymers: [Polymer Ambassadors]*** Use paper clips as monomers. Put two together in a chain to make a dimer, three in a chain for a trimer and keep increasing the length of the chain for a polymer. Use two different size paper clips to build a copolymer model. Place several of each type of paper clip together in the copolymer chain to represent a block polymer.

Use of Marshmallows:

1. ***Molecular Models:*** Use tooth picks and marshmallows to build models of molecules with the correct shapes, e.g. tetrahedral for methane CH_4, angular for water H_2O, etc.

2. ***Law of Conservation of Mass:*** Marshmallows are made to very exact masses so simulate a reaction with reactants and products represented with toothpicks and marshmallows on a double pan balance and show that the atoms and bonds must balance in a chemical reaction, hence the balanced chemical equation. A reaction represented by: $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$.

3. ***Indicators of a Chemical Change:*** Take a marshmallow on a skewer and roast it above a Bunsen burner flame. Note that the colour changes, the odour changes, the size changes, the taste changes, the reaction is difficult to reverse, and if heated hot enough visible heat and light [a flame] are produced. The only indicator of chemical change not present is the formation of a precipitate. This beats producing H_2S (g) as an odour change! [Note that the expansion of the marshmallow on heating can be explained with Charles’ Law.]

4. ***Boyle’s Law: [Pat Funk, David Katz]*** Place a marshmallow into a syringe. Carefully push as much air out of the syringe as possible. Lock the syringe. Pull the plunger out lowering internal pressure. What happens to the volume of the marshmallow? Why? Now place a marshmallow in a syringe. Place the plunger in. Close off the syringe and increase the pressure on the marshmallow. What happens this time? Why?

5. ***Nuclear Reactions: [AEC]*** Use one colour marshmallow for protons another colour for neutrons. Place the correct number of each in a large zip lock bag to represent a nucleus. For nuclear reactions make a nucleon for decay out of a marshmallow of the opposite colour and a small marshmallow of the correct colour held together with a toothpick. For example, if protons are white and neutrons red, use a large white marshmallow and a small red marshmallow. The red indicates neutron. On decay remove the small red marshmallow, an electron, which leaves a proton behind. For fusion, place two small zip lock bags into a larger bag and then follow correct decay.

Some Environmental Chemistry

Generate and collect the sulfur dioxide in a fume cupboard.

Purpose: to see the effects of acid deposition on materials

Materials: Erlenmeyer flasks, stoppers, gas delivery tube, 3.0 mol/L sulfuric acid, sodium sulfite, litmus paper, red flower petals, marble chips, iron nail, copper strip, 10 mL graduated cylinder, distilled water.

Procedure: 1. In a fume cupboard, place 30 g of sodium sulfite, add 20 ml of 3.0 mol/L sulfuric acid slowly through a thistle tube in a stopper to the solid sodium sulfite. Collect the sulfur dioxide gas produced in 250 mL Erlenmeyer flasks by the upward displacement of air. Stopper the flasks when full.

2. Into one flask place 5 mL of distilled water and restopper the flask. This is an acid rain flask. A second flask is the acid deposition or acid air flask.
3. Into each flask place a piece of red and blue litmus paper, a red flower petal, a marble chip, an iron nail, a piece of copper.
4. Observe each material before and after it enters the flask, and once it comes out of the flask. Describe the changes. Remove the materials in the fume cupboard.

The blue litmus will turn red, then the litmus and the flower petal will lose colour showing the bleaching power of sulfur dioxide.
The nail will rust, but will be a black colour, the copper may begin to give a blue or green colour indicating corrosion.
The changes occur faster in the wet flask
The marble chip looks the same but the outside will flake off when handled.

The procedure can be repeated with nitrogen oxides in the flasks if you can have concentrated nitric acid. Produce the nitrogen oxides, mainly NO_2, in a flask in a fume cupboard by adding 10 mL of concentrated nitric acid to 10 g of copper turnings. Collect the gas in Erlenmeyer flasks in the fume cupboard by the upward displacement of air. Add the same materials.

For a comparison, if you have a source of oxygen, collect flasks of oxygen and observe the effect of oxygen on the same materials. [The oxygen is nowhere near as corrosive.]

Safety Rules for the Labs

Most science safety rules are written for chemistry. However physics and biology have their own safety considerations. Here are some safety rules for physics and biology for you to take and adapt and edit and add to for those classes at you school.

SAFETY RULES: PHYSICS

Many people think that physics is a safe science, but doing experiments in physics does have dangers. Please follow the following rules when carrying out physics experiments.

1. When using optical equipment, lenses, mirrors, prisms, many are made of glass. The glass can chip and have sharp edges. These sharp edges can cut you. Please take care to check for any sharp edges before doing the experiment. Report the faulty equipment to the teacher and use safe materials. ***Wear safety goggles.***

2. The glass may chip while you are using it. ***Wear safety goggles*** until all equipment is put away.

3. Quite often candle flames are used in optics experiments. Keep hair out of the flame, do not burn yourself. Please clean up any wax that may drip off the candle.

4. When using electrical equipment, especially with power supplies, check for frayed wire, bare metal at connections, and report these to your teacher. Some metal pieces are sharp and can cut you. Again use safe equipment.

5. Because of electrical connectors, there may be bare metal exposed. Be aware of this and do not touch the bare metal. **Avoid shocks!**

6. When doing experiments with moving objects, be aware of the path the moving object will take. Warn other groups if they are in the way. Keep out of the way of the moving object. Do not be hit or have others hit with a projectile. ***Wear safety goggles, moving objects can cause eye injury!***

7. Beware of falling objects in any experiment where objects are dropped.

8. Beware of moving parts on any motorized equipment. Do not get your clothing or any body parts caught in the moving parts.

9. Keep people back from moving springs. Be sure the springs are secure. Do not overextend and damage the springs. Do not allow the springs to damage you.

Safety Rules: BIOLOGY

1. Use the microscopes with care. These are expensive instruments.

2. Most microscope slides are made of glass. If chipped the slides will be sharp and cut you. Handle the slides with care. Report chipped slides to your teacher.

3. Commercially prepared microscope slides are expensive. Some are $20.00 each. Handle these with care. Again they are glass and need to be checked for chipping so you are not cut.

4. Slides need to be placed cover slip side up on the microscope for proper focusing.

5. When using equipment for dissecting, note that you are using sharp utensils. Do not cut yourself.

6. Dissecting specimens are preserved in liquids. Be careful when putting in a first cut. The liquid may spray on you. Wear gloves and use friction tape on your fingers to minimize cuts. Wear protective clothing and ***eye protection.***

7. Dispose of any dissected specimens as instructed by your teacher.

8. When doing experiments requiring heating devices, do not get burned. Hot plates and burner heated equipment may not look hot, but will be hot enough to cause serious burns. Let things cool.

9. When using bacteria, mould or other live cultures, keep yourself and others from being contaminated. Some of these organisms can cause disease.

10. When using chemicals, please follow the chemistry safety rules.

11. When in doubt, ask the teacher for guidance.