



# Determination of the Charge on an Electron

## Introduction

Robert Millikan, an American scientist, determined the charge on a single electron in 1909. You will be collecting data to make the same determination of the charge on an electron. Although your procedure will be different than Millikan's, your calculations should yield the same charge. You will be collecting zinc atoms on a zinc electrode to carry out this activity. Zinc ions,  $\text{Zn}^{2+}$ , will each gain 2 electrons to become neutral zinc atoms that "electroplate" on one of the zinc electrodes. The other zinc electrode will lose the same number of zinc atoms as they become zinc ions.

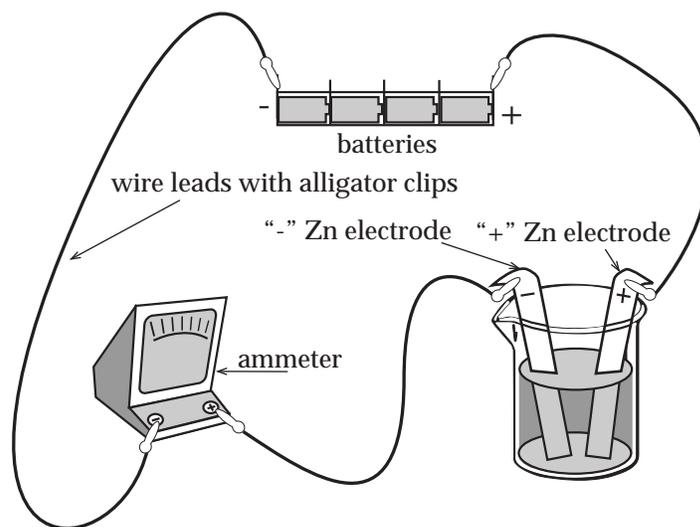
## Procedure

1. Put on goggles. Lay the zinc electrodes on paper towels on your lab counter and polish both electrodes on both sides by rubbing with steel wool. Set electrodes on clean counter and carefully wad up paper towels with steel wool residue and throw away.
2. Using a pencil, label one electrode as "+" at the top of the metal zinc strip and the other as "-".
3. Find the mass of the "+" electrode and record. Repeat with the "-" electrode.
4. Place the electrodes in a 250 mL beaker so that the labeled ends are at the top of the beaker and the electrodes are on opposite sides of the beaker. Bend the tops over the rim of the beaker.
5. Place the 4 batteries all facing the same direction in the battery holder. Connect the "+" end of the batteries to the "+" electrode with the alligator clips of a wire lead. In similar fashion connect the "-" battery end to the "-" terminal on the ammeter with another wire lead. Finally, connect the "+" ammeter terminal lead to the "-" electrode. Have the set-up checked by your teacher before you continue. (*see illustration to the right*)
6. Remove the alligator clip from the "-" ammeter terminal. Pour  $\text{ZnSO}_4$  solution into the 250 mL beaker containing the electrodes until it is about an inch or so from the rim.
7. Reconnect the ammeter and quickly adjust the current until the ammeter reads about 0.70 amperes. Record the amperage every minute for at least 20 minutes - 25 minutes is preferable!
8. Monitor the ammeter during data collection. If the reading changes 0.05 amps or more during a one-minute interval, wait until it is time to take the next reading and then readjust the current to about 0.70 amps.

(continued on back side)

## Materials

- 2 zinc electrodes (~3 X 10 cm)
- steel wool
- balance
- 250 mL beaker
- ~200 mL 1.0 M  $\text{ZnSO}_4$  solution
- 3 wire leads with alligator clips
- battery holder
- 4 size "D" batteries
- ammeter
- distilled  $\text{H}_2\text{O}$  in squirt bottle
- goggles
- clock or watch with second hand





# Determination of the Charge on an Electron

## Data Table

Initial mass of “+” electrode \_\_\_\_\_g

Initial mass of “-“ electrode \_\_\_\_\_g

| Time (Minutes) | Current (Amperes) |
|----------------|-------------------|
| 1              |                   |
| 2              |                   |
| 3              |                   |
| 4              |                   |
| 5              |                   |
| 6              |                   |
| 7              |                   |
| 8              |                   |
| 9              |                   |
| 10             |                   |
| 11             |                   |
| 12             |                   |
| 13             |                   |
| 14             |                   |
| 15             |                   |
| 16             |                   |
| 17             |                   |
| 18             |                   |
| 19             |                   |
| 20             |                   |
| 21             |                   |
| 22             |                   |
| 23             |                   |
| 24             |                   |
| 25             |                   |

Final mass of “+” electrode \_\_\_\_\_g

Final mass of “-“ electrode \_\_\_\_\_g

## Procedure *(continued from front side)*

- After your final timed ammeter reading, disconnect the wire leads. Carefully remove the electrodes and rinse with distilled H<sub>2</sub>O from squirt bottle over sink. Place on clean paper towel that has been labeled with your names. Allow electrodes to dry over night in a clean safe place such as a drawer or cupboard.
- Pour ZnSO<sub>4</sub> solution back into stock container so it can be reused. Clean up according to your teacher’s directions and wash your hands.
- After a drying period of at least 24 hours, find the mass of each electrode and record.

## Analysis and Conclusions

- Find the average of your ammeter readings.
  - Find the total time that you ran the experiment in seconds.
  - Multiply the average amperage by the total time in seconds to determine the total charge in ampere-seconds.
- Calculate the change in the mass of each electrode.  
 “+” electrode: \_\_\_\_\_g  
 “-“ electrode: \_\_\_\_\_g
  - Calculate the number of zinc atoms gained or lost at each electrode.  
 “+” electrode: \_\_\_\_\_ Zn atoms  
 “-“ electrode: \_\_\_\_\_ Zn atoms
  - Divide the total charge (1. c above) by the number of atoms at each electrode (3 above) to calculate the charge per atom.  
 “+” electrode: \_\_\_\_\_ [(ampere-seconds)/Zn atoms]  
 “-“ electrode: \_\_\_\_\_ [(ampere-seconds)/Zn atoms]
  - Calculate the charge per electron - remember that each Zn atom produces 2 electrons as it becomes a Zn<sup>2+</sup> ion!  
 \_\_\_\_\_ [(ampere-seconds)/electron]  
 \_\_\_\_\_ [(ampere-seconds)/electron]
  - If one ampere-second equals one coulomb, how do your calculations compare to the accepted value as determined by Millikan of 1.60 X 10<sup>-19</sup> coulombs? What is your % error for each electrode?  
 “+” electrode: \_\_\_\_\_ % error  
 “-“ electrode: \_\_\_\_\_ % error



# Determination of Moles of Copper and Iron in Reaction

## Introduction

In this activity you will perform a single replacement reaction between copper and iron. Using the masses of iron used and copper produced, you will be able to calculate the number of moles involved in the reaction.

## Procedure

### Day 1

1. Put on goggles. Over a paper towel, clean 2 iron nails with a small piece of sandpaper until they are shiny. Carefully wad up the paper towels with sandpaper residue and throw away. Find the mass of the 2 iron nails and record.
2. Label the 250 mL beaker with your initials using the wax pencil. Find the mass of the beaker and record.
3. Add about 2.0 g of copper(II) chloride and find the mass of the beaker with the copper(II) chloride and record.
4. Add about 50 mL distilled H<sub>2</sub>O to the beaker and gently swirl to dissolve the copper(II) chloride.
5. Carefully put the nails in the copper(II) chloride solution and then place the beaker in a safe location to allow the reaction to proceed overnight. Clean up as directed by your teacher and wash your hands.

### Day 2

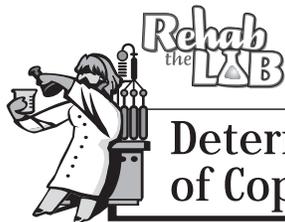
6. Put on goggles. Using tongs, carefully pick up the nails from the bottom of the beaker. With the squirt bottle of distilled H<sub>2</sub>O, rinse any copper clinging to the nails back into the beaker.
7. Place the nails on a paper towel labeled with your names. Allow the nails on the paper towel to dry in a safe place overnight. A drawer or cupboard will work well.
8. Decant the liquid in the 250 mL beaker into another large beaker. Use a stirring rod to help guide the liquid into the large beaker.
9. Add about 25 mL distilled H<sub>2</sub>O to the copper formed in the 250 mL beaker and swirl to rinse. Decant into large beaker. Repeat the rinsing/decanting procedure 3 more times. Dispose of the decant down the sink and flush with excess water.
10. Next add about 25 mL 1 M HCl to the copper and swirl. Decant into large beaker. Dispose of HCl decant as directed by your teacher.
11. Finally, add about 25 mL distilled H<sub>2</sub>O and swirl. Decant.
12. Place clean copper in labeled 250 mL beaker in oven to dry for several hours. Clean up as directed by your teacher and wash your hands.

### Day 3

13. Put on goggles. Find the mass of your 2 iron nails and record. Find the mass of the 250 mL beaker with dry copper and record. Clean up as directed by your teacher and wash your hands.

## Materials

- 2 iron nails
- sand paper
- balance
- 250 mL beaker
- wax pencil
- copper(II) chloride
- distilled H<sub>2</sub>O
- tongs
- squirt bottle of distilled H<sub>2</sub>O
- stirring rod
- large beaker for decanting
- 1 M HCl
- drying oven
- goggles



## Determination of Moles of Copper and Iron in Reaction

### Data

- (A) Initial mass of 2 iron nails: \_\_\_\_\_g
- (B) Mass of empty 250 mL beaker: \_\_\_\_\_g
- (C) Mass of 250 mL beaker plus copper(II) chloride: \_\_\_\_\_g
- (D) Final mass of 2 iron nails: \_\_\_\_\_g
- (E) Mass of 250 mL beaker plus dry copper: \_\_\_\_\_g

### Analysis and Calculations

1. Find the mass of iron used from the nails. \_\_\_\_\_g
2. Find the mass of copper produced in the 250 mL beaker. \_\_\_\_\_g
3. Find the moles of iron used from the nails. \_\_\_\_\_moles
4. Find the moles of copper produced. \_\_\_\_\_moles
5. What is the ratio of moles of iron used to moles of copper produced?
6. Write a balanced equation for this reaction.



# Determination of Density

## Introduction

Density is mass per unit of volume and is characteristic of a substance regardless of sample size. You will be calculating the densities of three different unknown liquids in this exercise by pipetting a known volume of each liquid and then measuring its respective mass. Using data generated by the entire class, you will then graph volume on the abscissa (x-axis) and mass on the ordinate (y-axis) for each of the three liquids. The slope of each line produced on the graph will give you density in g/mL.

## Procedure

1. Put on goggles. Find the mass of a clean dry beaker and record.
2. Using your assigned volumetric pipet and pipet filler, measure that volume of one of the unknown liquids. NEVER MOUTH PIPET! (*see illustration to the right*)
3. Deliver the liquid to the beaker and find the mass of the beaker and the liquid. Record.
4. Pour the liquid back into the dispensing container. Thoroughly wash and dry the beaker. Wash the pipet with distilled H<sub>2</sub>O and carefully shake dry.
5. Repeat the procedure using the other unknown liquids.
6. Clean up as directed by your teacher and wash your hands.

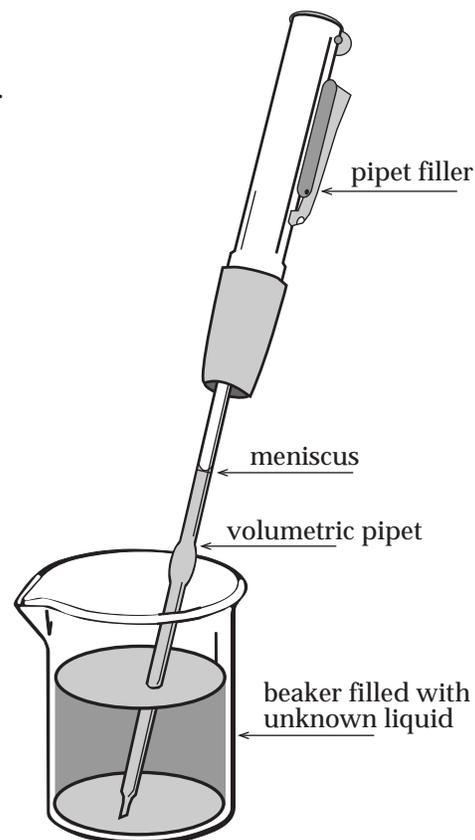
## Materials

- small beaker (25 or 50 mL)
- centigram balance
- volumetric pipet
- goggles
- pipet filler
- graph paper
- liquid A
- liquid B
- liquid C

## Data Table

Volumetric pipet size assigned to you: ml

|                             | Liquid A | Liquid B | Liquid C |
|-----------------------------|----------|----------|----------|
| Mass of small beaker        | g        | g        | g        |
| Mass of beaker and liquid   | g        | g        | g        |
| Mass of liquid (subtract)   | g        | g        | g        |
| Density (mass/pipet volume) | g/mL     | g/mL     | g/mL     |





## Determination of Density

### Class Data

|    | Partners' Initials | Volume | Mass A | Mass B | Mass C |
|----|--------------------|--------|--------|--------|--------|
| 1  |                    |        |        |        |        |
| 2  |                    |        |        |        |        |
| 3  |                    |        |        |        |        |
| 4  |                    |        |        |        |        |
| 5  |                    |        |        |        |        |
| 6  |                    |        |        |        |        |
| 7  |                    |        |        |        |        |
| 8  |                    |        |        |        |        |
| 9  |                    |        |        |        |        |
| 10 |                    |        |        |        |        |
| 11 |                    |        |        |        |        |
| 12 |                    |        |        |        |        |
| 13 |                    |        |        |        |        |
| 14 |                    |        |        |        |        |
| 15 |                    |        |        |        |        |
| 16 |                    |        |        |        |        |

### Analysis and Conclusions

After you have collected data from each pair of students in your class, choose three colored pencils to represent each of the three liquids. Label the abscissa as “volume” and the ordinate as “mass” and write in numbers to represent the range for class data for both mass and volume. Plot points for the three liquids in the respective colors you have chosen and draw best fit lines. Calculate the slope (rise/run) for each line and you will have determined the density of each liquid.

Slope of liquid A line = \_\_\_\_\_

Slope of liquid B line = \_\_\_\_\_

Slope of liquid C line = \_\_\_\_\_



# Determination of the Molar Volume of a Gas

## Introduction

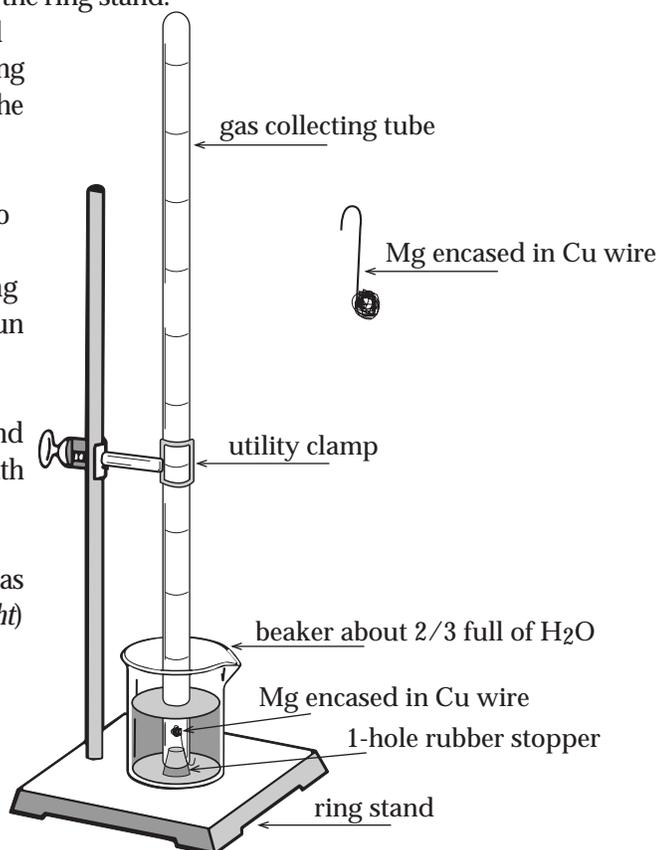
Equal volumes of gases contain equal numbers of gaseous particles. So how much volume would  $6.02 \times 10^{23}$  gaseous particles occupy at standard temperature and pressure? Using a known amount of Mg in reaction with HCl, you will be able to make this determination.

## Procedure

1. Put on goggles. Using a ruler, measure the length of your magnesium ribbon to the nearest 0.1 cm and record.
2. Roll the Mg into a small ball and wrap a piece of fine Cu wire (~ 25 cm or so) around the rolled Mg to encase it. As you wrap the Mg, leave about 4-5 cm of Cu wire extended from the "case" as a "tail." Bend the last cm of the Cu wire tail over to form a hook. See illustration.
3. Place the utility clamp on the ring stand. Fill the large beaker about 2/3 full of room temperature tap H<sub>2</sub>O. Use your thermometer and adjust the hot and cold H<sub>2</sub>O faucets as needed.
4. Record room temperature and barometric pressure.
5. Secure gas measuring tube in utility clamp. The open end should be on top and the bottom should be about 15 cm above the base of the ring stand.
6. Measure 10 mL of 6.0 M HCl in the graduated cylinder and carefully pour into the gas measuring tube. Without disturbing the HCl, slowly pour H<sub>2</sub>O into the gas measuring tube until the H<sub>2</sub>O comes to the brim of the tube. Place the beaker of H<sub>2</sub>O under the gas measuring tube on the base of the ring stand.
7. Hook the Cu wire over the top of the gas measuring tube so that the encased Mg is in the H<sub>2</sub>O in the tube. Insert the 1-hole rubber stopper securely in the top of the gas measuring tube. A little H<sub>2</sub>O should be displaced by the stopper and run down the outside of the tube into the beaker underneath.
8. Cover the hole in the stopper with your finger while your partner loosens the utility clamp. Quickly invert the tube and place the stoppered end of the gas measuring tube underneath the H<sub>2</sub>O level in the beaker. Secure the tube in the utility clamp with the stopper about 1/2" above bottom of beaker. Observe the dense HCl move down through the H<sub>2</sub>O in the gas measuring tube to react with the Mg. (see illustration to the right) (continued on back side)

## Materials

- gas measuring tube
- ring stand
- utility clamp
- large beaker (400 or 600 mL)
- graduated cylinder (10 mL)
- thermometer
- barometer
- 1-hole rubber stopper (to fit gas measuring tube)
- large deep container of H<sub>2</sub>O
- ruler
- magnesium ribbon
- fine copper wire
- 6.0 M HCl





## Determination of the Molar Volume of a Gas

### Procedure *(continued from front side)*

9. Wait about 5 minutes after the reaction has stopped and then tap the tube to dislodge any bubbles that may have formed.
10. Cover the hole in the stopper with your finger while your partner loosens the utility clamp. Quickly walk (stopper end down, but sealed with your finger!) to the large container of  $\text{H}_2\text{O}$ . Place the stoppered bottom of the tube under the  $\text{H}_2\text{O}$  level in the large container. Remove your finger that is sealing the stopper. Raise or lower the tube until the  $\text{H}_2\text{O}$  level in the tube is the same as the  $\text{H}_2\text{O}$  level in the container. Record the volume of collected gas.
11. Reseal the stopper with your finger, remove from the large container, and quickly turn the gas measuring tube right side up. Dispose of the excess  $\text{HCl}$  and clean up as directed by your teacher. Wash your hands.

### Data Table

Mass of 1.00 m of Mg ribbon (given by teacher) \_\_\_\_\_ g  
Length of Mg ribbon \_\_\_\_\_ g  
Room Temperature \_\_\_\_\_ °C  
Barometric Pressure \_\_\_\_\_ mm Hg  
Volume of Collected Gas \_\_\_\_\_ mL

### Analysis and Calculations

1. Write the balanced equation for the reaction.
2. Calculate the mass of Mg used.
3. Calculate the moles of Mg used.
4. Using your balanced equation, determine how many moles of  $\text{H}_2$  you produced.
5. Convert barometric pressure from mm Hg to kPa.
6. Using Dalton's Law, find the partial pressure of the  $\text{H}_2$  collected over  $\text{H}_2\text{O}$ .
7. Using the combination of Boyle's and Charles' gas laws, determine what your volume of  $\text{H}_2$  would be if collected at standard temperature ( $273^\circ \text{K}$ ) and standard pressure (101.3 kPa).
8. Convert your standard gas volume to liters.
9. You produced far less than a mole of  $\text{H}_2$ . Using a ratio proportion, calculate the volume that 1 mole of gas would occupy at standard temperature and pressure.



# Formula of a Hydrate

## Introduction

Many ionic compounds incorporate a fixed number of water molecules into their crystal structures. These are called hydrates. Heat can be applied to a hydrated salt to release the  $H_2O$  molecules and produce an anhydrous salt which often will appear different than its hydrate. When expressing the formula for a hydrate, it is necessary to notate the fixed number of  $H_2O$  molecules following the anhydrous formula for the ionic compound. A large dot is placed between the formula and the  $H_2O$  molecules. For example:  $CuSO_4 \cdot 5H_2O$  is the correct formula for the hydrated form of copper sulfate. In this activity, you will be determining the number of  $H_2O$  molecules in the hydrate of either alum (aluminum potassium sulfate:  $AlK(SO_4)_2$ ) or Epsom salts (magnesium sulfate:  $MgSO_4$ ).

## Procedure

1. Put on your goggles. Secure iron ring on ring stand a couple of inches above the height of the burner. Place wire gauze (alum) or triangle (Epsom salts) on iron ring.
2. Place clean evaporating dish or crucible and cover on gauze or triangle. Light burner and heat for a couple of minutes to make certain container is thoroughly dry. Turn off burner and cool container for several minutes until it is comfortable to touch. Record the mass of the dry evaporating dish (alum) or crucible and cover (Epsom salts).
3. Add about a tablespoon of alum to the evaporating dish if that is the hydrate assigned to you. If your assigned hydrate is Epsom salts, add about  $\frac{1}{2}$  teaspoon and cover the crucible. Record the mass of the container with the hydrate.
4. Place the container back on the gauze/triangle (cover should be slightly ajar) and heat gently with hot flame until the water has been released from the hydrate. This will require about 5 minutes. (see illustrations below)

(continued on back side)

## Materials

(Alum)

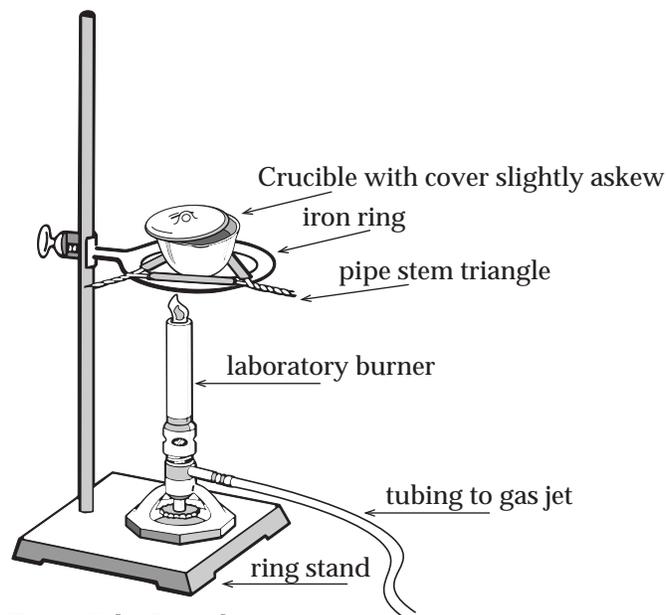
- evaporating dish
- wire gauze
- alum ~1 tablespoon

(Epsom salts)

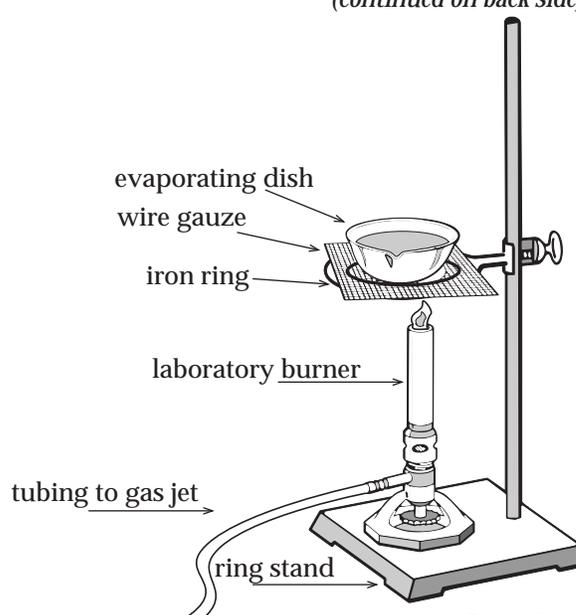
- crucible and cover
- pipe stem triangle
- Epsom salts ~1/2 teaspoon

(Both)

- crucible tongs
- balance
- ring stand
- iron ring
- laboratory burner
- burner lighter
- goggles



Epsom Salts Procedure



Alum Procedure



## Formula of a Hydrate

### Procedure *(continued from front side)*

- When no more  $\text{H}_2\text{O}$  appears to be coming from the hydrate, turn off the burner and cool for several minutes until container is comfortable to the touch.
- Record the mass of the container (and cover if using Epsom salts) with anhydrous salt.
- If time allows, reheat the container with salt, cool and remass. If the two final masses agree, you can be confident that you have indeed released all of the  $\text{H}_2\text{O}$  from the hydrate.
- Clean up as directed by your teacher and wash your hands.

### Data Table

Hydrate assigned to you:

Alum

|    |                                 |   |
|----|---------------------------------|---|
| a. | Mass of Evaporating Dish        | g |
| b. | Mass of Dish and Hydrate        | g |
| c. | Mass of Dish and Anhydrous Salt | g |

Epsom Salts

|    |  |   |
|----|--|---|
| a. | Mass of Crucible and Cover               | g |
| b. | Mass of Crucible, Cover, & Hydrate       | g |
| c. | Mass of Crucible, Cover & Anhydrous Salt | g |

### Analysis and Calculations

- Calculate the mass of the anhydrous salt:

\_\_\_\_\_g

- Find the molar mass of anhydrous alum ( $\text{Al}_2(\text{SO}_4)_3$ ) or Epsom salts ( $\text{MgSO}_4$ ):

\_\_\_\_\_g

- Calculate the moles of anhydrous salt:

\_\_\_\_\_moles

- Calculate the mass of  $\text{H}_2\text{O}$  "cooked out" of your hydrate:

\_\_\_\_\_g

- Find the molar mass of  $\text{H}_2\text{O}$ :

\_\_\_\_\_g

- Calculate the moles of  $\text{H}_2\text{O}$  released:

\_\_\_\_\_moles

- Divide the moles  $\text{H}_2\text{O}$  by the moles of anhydrous salt to determine the ratio of moles of  $\text{H}_2\text{O}$  to moles of anhydrous salt. Round to the nearest small whole number.

\_\_\_\_\_

- Write the correct formula for your hydrate: \_\_\_\_\_ • \_\_\_\_\_  $\text{H}_2\text{O}$ .

- Using the correct formula, calculate the percent of  $\text{H}_2\text{O}$  in your hydrate.



# Heating and Cooling Curves

## Introduction

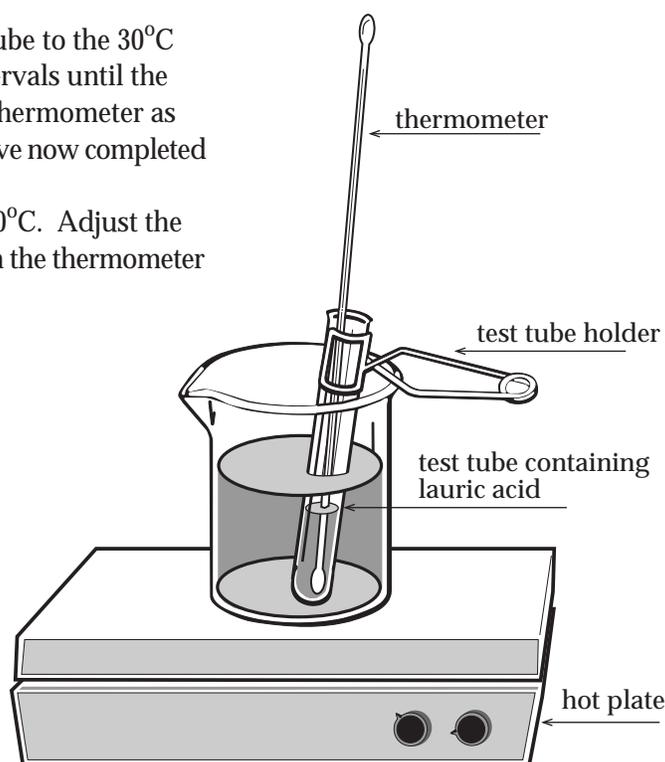
If you were to record the temperature of a pure solid substance for a period of time as you heated the substance beyond its melting point, what would the graph of temperature vs. time look like? How about a graph generated from temperatures taken during cooling? In this exercise, you will collect data on the heating and cooling of pure lauric acid.

## Procedure

1. Put on goggles. Fill one beaker about 2/3 full of warm tap water and place on hot plate. Adjust the hot plate so that you can maintain the temperature of the water at about 60°C.
2. Adjust the hot/cold water faucets so that the running water is about 30°C. Fill the other beaker about 2/3 full of 30°C water from the tap and set aside.
3. Place the test tube of lauric acid in the 60°C water on the hot plate. When the lauric acid starts to melt, carefully put a thermometer down into the melting solid. Continue to heat the lauric acid until it is about 55°C. (*see illustration to the right*)
4. Using a test tube holder, transfer the liquid lauric acid tube to the 30°C water bath and record the temperature at 30-second intervals until the temperature reaches at least 40°C. Gently stir with the thermometer as you take data until the lauric acid starts to solidify. You have now completed the cooling curve data.
5. Check your 60°C water bath to be sure that it is still at 60°C. Adjust the hot plate if necessary and transfer the lauric acid tube with the thermometer back into the warm water.
6. Record the temperature of the lauric acid at 30-second intervals until the temperature reaches about 50°C. As soon as it is melted enough to stir the lauric acid with a thermometer, carefully do so. You now have completed the heating curve data.
7. Quickly remove the thermometer from the lauric acid and thoroughly wipe the lauric acid off with a tissue (Kim-Wipe®). Return the lauric acid in the test tube to dispensing area.
8. Clean up as directed by your teacher and wash your hands.

## Materials

- large test tube containing about 15 g lauric acid
- test tube holder
- 2 large (400 or 600 mL) beakers
- hotplate
- 2 thermometers
- clock or watch with second hand
- goggles
- graph paper





## Heating and Cooling Curves

### Data Table

| Time (min.) | Cooling          | Heating          |
|-------------|------------------|------------------|
|             | Temperature (°C) | Temperature (°C) |
| 0.5         |                  |                  |
| 1.0         |                  |                  |
| 1.5         |                  |                  |
| 2.0         |                  |                  |
| 2.5         |                  |                  |
| 3.0         |                  |                  |
| 3.5         |                  |                  |
| 4.0         |                  |                  |
| 4.5         |                  |                  |
| 5.0         |                  |                  |
| 5.5         |                  |                  |
| 6.0         |                  |                  |
| 6.5         |                  |                  |
| 7.0         |                  |                  |
| 7.5         |                  |                  |
| 8.0         |                  |                  |
| 8.5         |                  |                  |
| 9.0         |                  |                  |
| 9.5         |                  |                  |
| 10.0        |                  |                  |
| 10.5        |                  |                  |
| 11.0        |                  |                  |
| 11.5        |                  |                  |
| 12.0        |                  |                  |
| 12.5        |                  |                  |
| 13.0        |                  |                  |
| 13.5        |                  |                  |
| 14.0        |                  |                  |
| 14.5        |                  |                  |
| 15.0        |                  |                  |

### Graphing

Choose a colored pencil for your cooling data and another colored pencil to represent your heating data. Label the abscissa (X-axis) on your graph paper as “time” and ordinate (y-axis) as “temperature” and add numbers to represent the ranges for your data. Plot data in the colors you have chosen.

### Analysis and Conclusions

1. What is the freezing point for lauric acid? What is the melting point for lauric acid?
2. Explain what is happening during each section of the graph for both cooling and heating in terms of particle motion and energy.



# I Scream, You Scream, We all Scream For Ice Cream!

## Introduction

How do old-fashioned homemade ice cream makers work? You will investigate the answer to that question as well as enjoy some homemade ice cream in this activity.

## Procedure

1. Put on goggles. Make certain your work area is thoroughly clean. Fill the pop bottle bottom almost full of ice chips and generously sprinkle rock salt on top.
2. Pour ice cream mix into plastic cup until it is about 2 inches deep. This should be about  $\frac{1}{2}$  cup of ice cream mix.
3. Carefully work the bottom of the cup of the ice cream mix down into the ice/rock salt mixture. The top of the ice cream mix should be at or below the level of the ice.
4. Stir occasionally with a spoon. As ice cream starts to solidify, you may wish to stir more frequently. Add rock salt to ice as needed, but be careful not to get salt in your ice cream!
5. Just before ice cream looks ready, record the temperature of the ice/rock salt mixture, but do not let the thermometer touch the ice cream.
6. When ice cream is the consistency of soft serve, it is done. Remove cup from ice/rock salt and ENJOY!
7. Clean up as directed by your teacher and wash your hands.

## Data Table

Temperature of Ice/Rock Salt Mixture: \_\_\_\_\_°C

## Analysis and Calculations

1. You probably used about 500 g of ice. How many grams of rock salt were dissolved in it?
2. How can you relate making homemade ice cream to salting sidewalks and roads during the winter?
3. How do the new ice cream makers (like Donvier<sup>®</sup>) work?

## Materials

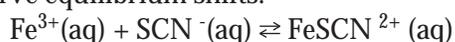
- bottom half of 2 L pop bottle
- plastic cup
- spoon
- ice
- rock salt
- ice cream mix
- thermometer



# Investigating Le Chatelier's Principle

## Introduction

Le Chatelier's principle states that if you have a system at equilibrium and then change something in the system, the equilibrium will shift to counteract the change. In this investigation, you will use the following reversible reaction to observe equilibrium shifts:



The ions involved in this system have characteristic colors that will allow you to follow shifts in equilibrium:  $\text{Fe}^{3+}$  is gold,  $\text{SCN}^{-}$  is colorless and  $\text{FeSCN}^{2+}$  is orange-rust.

## Procedure

1. Put on goggles. Put 3 drops of 0.2 M  $\text{FeCl}_3$  in the 250 mL beaker. Add 3 drops of 0.2 M  $\text{KSCN}$ . Swirl to mix.
2. Add distilled  $\text{H}_2\text{O}$  to the beaker to dilute the dark color to light orange. This may require 40-50 mL of  $\text{H}_2\text{O}$  or more. Swirl to mix.
3. Label 5 test tubes A to E. Fill test tubes about  $\frac{3}{4}$  full of the light orange solution and place in test tube rack.
4. Test tube A will serve as a control.
5. Add 4 drops of 0.2 M  $\text{Fe}(\text{NO}_3)_3$  to test tube B.
6. Add 4 drops of 0.2 M  $\text{KCl}$  to test tube C.
7. Add 4 drops of 0.2 M  $\text{KSCN}$  to test tube D.
8. Add 4 drops of 6.0 M  $\text{NaOH}$  to test tube E. CAUTION: The  $\text{NaOH}$  is very caustic – be extremely careful with it!
9. Examine test tubes B – E in comparison to test tube A. Note your observations/descriptions in the data table.
10. Clean up as directed by your teacher and wash your hands.

## Materials

- Dropper bottles of:
  - 0.2 M  $\text{FeCl}_3$
  - 0.2 M  $\text{KSCN}$
  - 0.2 M  $\text{Fe}(\text{NO}_3)_3$
  - 0.2 M  $\text{KCl}$
  - 6.0 M  $\text{NaOH}$
- distilled  $\text{H}_2\text{O}$
- 250 mL beaker
- 5 test tubes
- test tube rack
- wax pencil



## Investigating Le Chatelier's Principle

### Data Table

| Chemical Added                | Observations/<br>Description | Stress Ion | Spectator Ion | Direction of<br>Equilibrium Shift |
|-------------------------------|------------------------------|------------|---------------|-----------------------------------|
| B: $\text{Fe}(\text{NO}_3)_3$ |                              |            |               |                                   |
| C: $\text{KCl}$               |                              |            |               |                                   |
| D: $\text{KSCN}$              |                              |            |               |                                   |
| E: $\text{NaOH}$              |                              |            |               |                                   |

### Analysis and Calculations

1. Using the equation in the introduction, determine the ion creating the stress in each test tube that had an equilibrium shift. Add to the data table.
2. Identify the spectator ion(s) for each test tube and add to data table.
3. Decide whether or not an equilibrium shift occurred in each test tube. On the data table, indicate the direction of equilibrium shifts with arrows.
4. How can you explain your observations in test tube E?



# Molar Mass Determination From Freezing Point Depression

## Introduction

Any solute added to a solvent will lower the freezing point of that solvent. Because the amount of solute in the solution determines how much the freezing point will be lowered and this phenomenon is independent of what the solute is, freezing point depression is known as a colligative property. Other colligative properties are boiling point elevation, vapor pressure reduction and osmotic pressure. The decrease in freezing point of the pure solvent is proportional to the molality of the particles of solute and is represented by the following equation:

$$\Delta T_f = K_f \cdot m$$

Where:

$\Delta T_f$  = decrease from freezing point of pure solvent

$K_f$  = freezing point depression constant (from a handbook)

$m$  = molality = (moles solute particles/kg solvent)

= [(grams solute particles/molar mass solute)/kg solvent]

In this lab, you will be measuring the freezing point of a solution of benzoic acid in lauric acid. Since you have already determined the freezing point of pure lauric acid in a previous lab, the only temperature measurements you will need to make will be to determine the freezing point of the benzoic acid/lauric acid solution. Your premeasured solution test tube contains 2.00 g of benzoic acid ( $C_6H_5COOH$ ) and 16.00 g of lauric acid.

## Materials

- large test tube containing 2.00 g benzoic acid + 16.00 g lauric acid
- test tube holder
- 2 large (400 or 600 mL) beakers
- hot plate
- 2 thermometers (one must be the same one used in the pure lauric acid for "Heating and Cooling Curves")
- clock or watch with second hand
- goggles
- graph paper

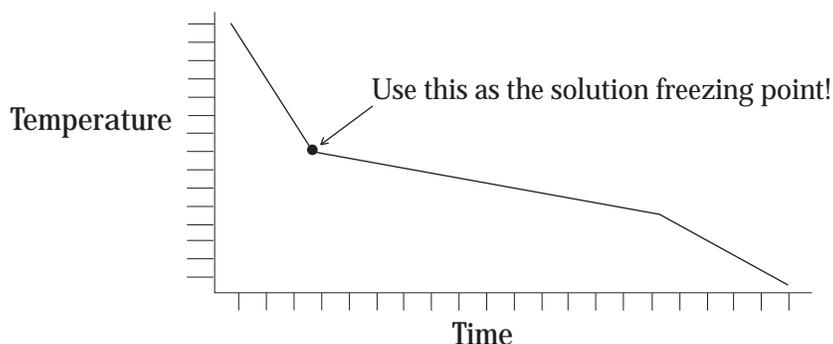




# Molar Mass Determination From Freezing Point Depression

## Graphing

Label the abscissa as “time” and the ordinate as “temperature” and add numbers to represent the ranges of your data. Plot the cooling curve data for the solution and determine the freezing point from the graph. Since solutions do not typically yield perfectly horizontal plateaus, choose the point at which the cooling rate drastically drops.



## Analysis and Calculations:

1. Subtract the solution freezing point from the pure lauric acid freezing point to obtain  $\Delta T_f$ .

$$\Delta T_f = \text{_____} \text{ } ^\circ\text{C}$$

2. Using  $\Delta T_f = K_f \cdot m$ , calculate the molality of the solution. The freezing point constant for lauric acid is  $3.9^\circ\text{C} - \text{kg}/\text{Mole}$ .

$$m = \text{_____} \text{ moles}/\text{kg}$$

3. Convert mass of lauric acid from g to kg.

$$\text{lauric acid} = \text{_____} \text{ kg}$$

4. Calculate moles of benzoic acid using:  $m = \text{moles solute}/\text{kg solvent}$ .

$$\text{benzoic acid} = \text{_____} \text{ moles}$$

5. Calculate molar mass of benzoic acid using:  $\text{moles} = \text{g}/\text{molar mass}$ .

$$\text{molar mass} = \text{_____} \text{ g}$$

6. Using the benzoic acid formula,  $\text{C}_6\text{H}_5\text{COOH}$ , and your periodic table, calculate the molar mass of benzoic acid.  
\_\_\_\_\_  $\text{g}/\text{moles}$

7. Determine your % error.

$$\text{_____} \%$$



# Percent Composition of an Egg

## Introduction

An egg is composed of three major parts: white, yolk, and shell. What percentage of the total by mass is each of these components? If you can solve this problem for an egg, you can calculate the percentage of each element in a compound if you know the chemical formula.

## Procedure

1. With your partner, devise your own plan to determine what percentage of your egg is white, yolk, and shell. After your plans are made, put on goggles and carry out your plan.
2. Dispose of all egg parts and clean up as directed by your teacher. Wash your hands thoroughly with soap.

## Data Table

|                                     | White | Yolk | Shell |
|-------------------------------------|-------|------|-------|
| Mass of Empty Container             |       |      |       |
| Mass of Container and Egg Component |       |      |       |
| Mass of Egg Component               |       |      |       |
| % of Total Egg                      |       |      |       |

## Materials

- egg
- assorted containers (beakers)
- balance
- hot plate (if you wish to cook the egg)
- goggles



# Solubility and Precipitation

## Introduction

Some combinations of ionic solutions produce insoluble products called precipitates and some produce no visible reaction at all. In this investigation you will be performing several double replacement reactions by combining several pairs of dissolved ionic substances. Using a solubility chart, you should be able to identify the precipitates that you observe. You also should be able to recognize some trends in solubility of some ions.

## Procedure

1. Put on goggles. Using your data table as a guide, combine 1 drop of the solution shown at the top of column with 1 drop of the solution shown to the left of the row in a well. For example, the well in the upper left-hand corner, (#1), should contain one drop of  $\text{AgNO}_3$  plus 1 drop of  $\text{CuSO}_4$ .
2. After you have made all 15 combinations, examine your well plate against a light background as well as a dark background. If the solution combination remains clear, record NR (no reaction) in the data table. If haziness or cloudiness occurs, a precipitate has formed. Record as PPT and also indicate the color of the precipitate.
3. Clean up as directed by your teacher and wash your hands.

## Materials

- 0.1 M solutions of the following chemicals (dispensed in labeled Beral pipets):  $\text{AgNO}_3$ ,  $\text{NaNO}_3$ ,  $\text{Na}_3\text{PO}_4$ ,  $\text{K}_2\text{CO}_3$ ,  $\text{FeCl}_3$ , and  $\text{CuSO}_4$ .
- well plate
- goggles

## Data Table

|                          | $\text{AgNO}_3$ | $\text{NaNO}_3$ | $\text{Na}_3\text{PO}_4$ | $\text{K}_2\text{CO}_3$ | $\text{FeCl}_3$ |
|--------------------------|-----------------|-----------------|--------------------------|-------------------------|-----------------|
| $\text{CuSO}_4$          | (1)             | (2)             | (3)                      | (4)                     | (5)             |
| $\text{FeCl}_3$          | (6)             | (7)             | (8)                      | (9)                     |                 |
| $\text{K}_2\text{CO}_3$  | (10)            | (11)            | (12)                     |                         |                 |
| $\text{Na}_3\text{PO}_4$ | (13)            | (14)            |                          |                         |                 |
| $\text{NaNO}_3$          | (15)            |                 |                          |                         |                 |



## Solubility and Precipitation

### Analysis and Conclusions

1. Write balanced equations for each double replacement reaction. If no precipitate formed, indicate the product as no reaction (NR). If a precipitate did form, consult your solubility chart to identify it. Indicate the insoluble product as a solid (s), and the remaining spectator ions as aqueous (aq).

(1)

(2)

(3)

(4)

(5)

(6)

(7)

(8)

(9)

(10)

(11)

(12)

(13)

(14)

(15)

2. Study your data for trends.

(A.) Which negative ion(s) did not form any precipitate?

(B.) Which positive ion(s) did not form any precipitate?

(C.) Which negative ion(s) usually or always formed a precipitate?

(D.) Which positive ion(s) usually or always formed a precipitate?