

# APPENDICES

---

Topic 1: Physical Properties of Matter Appendices

Topic 2: Gases and the Atmosphere Appendices

Topic 3: Chemical Reactions Appendices

Topic 4: Solutions Appendices

Topic 5: Organic Chemistry Appendices

General Appendices



# TOPIC 1:

# PHYSICAL PROPERTIES OF MATTER

## APPENDICES

---

Appendix 1.1: Demonstrating Diffusion on an Overhead Projector	3
Appendix 1.2: A “Real” Water Fountain	5
Appendix 1.3: Popping the Kernel: Modelling the States of Matter	6
Appendix 1.4: Kinetic Energy Distribution	10
Appendix 1.5: Bond Types and Conductivity	11
Appendix 1.6: Chemistry Article Review Form	15
Appendix 1.7: Making Fingerprints Visible	16
Appendix 1.8: Probeware Investigation: Determining Melting Points	17
Appendix 1.9: Vapour Pressure with Pop	22
Appendix 1.10: Measuring the Vapour Pressure of a Liquid	24
Appendix 1.11: Forces between Particles	27
Appendix 1.12: Freezing by Boiling	31
Appendix 1.13: Gas Laws: Temperature and Pressure Changes	32
Appendix 1.14: Chemistry Is Super: “Bingo” Review Game	36

**Note:**

Due to copyright considerations, Appendix 1.3: Popping the Kernel: Modelling the States of Matter is available only in the print document.



## Appendix 1.1: Demonstrating Diffusion on an Overhead Projector

### Introduction

A small amount of a liquid is placed in one compartment of a partitioned petri dish and two solids are sprinkled into two adjacent compartments. After a minute or so, some of the liquid has wet one of the solids and one of the solids has coloured the liquid.

### Chemical Concepts

- vapour pressure
- sublimation
- Raoult's Law
- diffusion
- equilibrium

### Materials

- four-quadrant glass petri dish with cover
- overhead projector and screen
- acetone
- paradichlorobenzene (or other relatively high vapour pressure solid)
- iodine

### Procedure

Sprinkle a few small crystals of paradichlorobenzene into one compartment of a four-quadrant petri dish and sprinkle a few crystals of iodine into the opposite compartment. Place the dish on an overhead projector stage. Add 1 or 2 mL of acetone to one of the remaining wells of the petri dish and cover it.

Have students observe the petri dish on the overhead projector. After a minute or so, small droplets of liquid can be observed surrounding the paradichlorobenzene crystals but not anywhere else in the dish. After a few minutes, the paradichlorobenzene crystals are completely dissolved in acetone and only droplets of solution appear. Some iodine has migrated into the edge of the acetone and is itself wet with acetone.

---

**Demonstrating Diffusion on an Overhead Projector:** Copyright © by Walter Rohr, Eastchester High School, Eastchester, NY. Adapted by permission of the author. All rights reserved.

**Appendix 1.1: Demonstrating Diffusion on an Overhead Projector** *(continued)*

---

**Discussion**

When the dish is sealed, migrating acetone molecules eventually collide with the surface of the paradichlorobenzene, forming bonds and, consequently, a solution. Since the vapour pressure of the paradichlorobenzene is lower than that of pure acetone (2 torr vs. 230 torr at 25°C), the resulting solution will have a lower vapour pressure than the pure acetone (Raoult's Law), preventing equilibrium from being achieved. As a result, higher vapour pressure acetone molecules continue to be forced into the lower vapour pressure solution in an attempt to produce equilibrium. If the demonstration is allowed to continue, the process will proceed until all the acetone has transferred into the paradichlorobenzene and iodine compartments. Eventually, the iodine can be observed discolouring the acetone at the edge nearest its compartment and itself being wet with the acetone.

## Appendix 1.2: A “Real” Water Fountain

### Introduction

A small quantity of water is heated in a 1 L flask to replace all the air. When the flask is filled with water vapour or steam, the heating is stopped. Steam condenses and the pressure inside the flask drops. Cold water from a beaker flows rapidly into the evacuated flask.

### Materials

- 7 mm glass tubing (bent)
- one-hole #8 rubber stopper
- Florence flask (1 L)
- beaker (1 L)

### Procedure

1. Fill the beaker with water.
2. Place 20 mL of water in the flask. Stopper the flask and use the tubing to connect the flask to the water-filled beaker.
3. Heat the flask. Gas from the warming flask will begin to flow into the beaker. Bubbles indicate the air being displaced from the flask.
4. Continue heating until the bubbling in the beaker ceases. This will happen by the time the water in the flask is at a full boil. Bubbles are no longer seen in the beaker if water vapour leaving the flask and entering the beaker is condensing.
5. Remove the source of heat from the flask. As the flask starts to cool, the water from the beaker will begin to flow from the beaker through the tube and into the flask. It can be spectacular!

### Safety Precaution

By using water, this demonstration is relatively safe, provided the tubing is carefully shaped and free of obstructions.

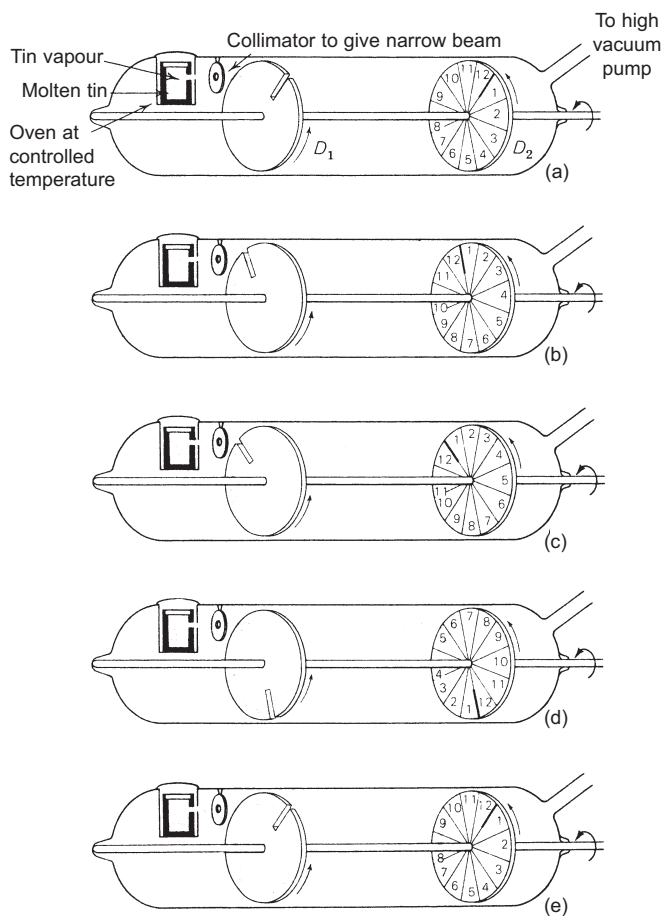
### Discussion

Steam from the heating water drives out the air from the flask. As the flask cools, the steam condenses and pressure inside the flask drops rapidly. Water from the beaker is then pushed into the flask by atmospheric pressure. Observe not only the powerful jet of water from the glass tube, but also the resurgence of boiling for any water remaining in the flask as the inside pressure is lowered. This boiling occurs almost simultaneously with the rushing of water into the flask.

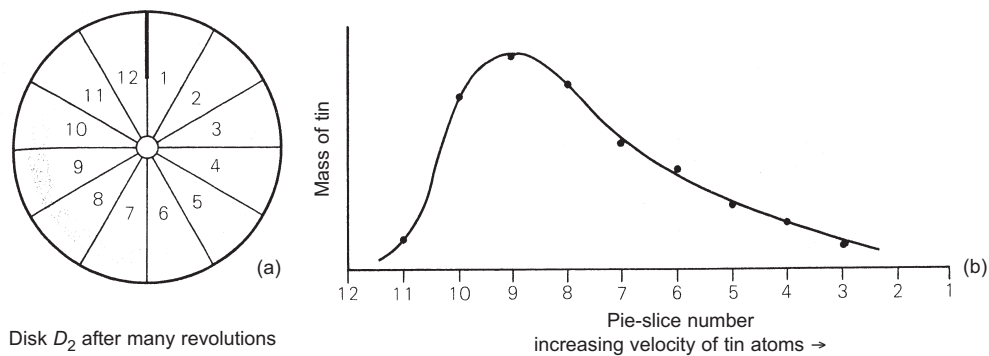
---

**A “Real” Water Fountain:** Copyright © by Penney Sconzo, Westminster High School, Atlanta, GA. Adapted by permission of the author. All rights reserved.

## Appendix 1.4: Kinetic Energy Distribution



Rotating disk used to measure atomic or molecular velocities.



The relative distribution of atomic or molecular velocities obtained from the rotating disk.

**Kinetic Energy Distribution Figures:** From *Chemistry: Experimental Foundations* Robert W. Parry, Phyllis M. Dietz, Robert L. Tellefsen & Luke E. Steiner © 1975 by Prentice Hall. Used by permission of Pearson Education, Inc.

## Appendix 1.5: Bond Types and Conductivity

### Question

What types of liquids or solutions conduct electricity and which ones do not?

### Prediction

Can a conductivity test determine the type of bond present within a substance?

### Materials

- computer system and conductivity sensor
- distilled water
- tap water
- ethyl alcohol,  $C_2H_5OH$
- 0.05 M concentration solutions of sugar ( $C_{12}H_{22}O_{11}$ ), table salt (NaCl), NaI, KCl, and KI
- 8 small beakers (100 mL)

### Safety Precautions

- Wear a lab apron and safety goggles.
- Ethyl alcohol is flammable. Keep away from flames.
- Table salt, NaCl, may cause eye irritation and pain.
- NaI is harmful if swallowed.
- KI is harmful if swallowed.
- Use caution with all solutions, as some are caustic, or poisonous, or will stain clothing.

### Procedure

1. Prepare a recording table as follows:

Data Table				
Trial #	Liquid or Solution	Predicted Type of Bond	Conductivity ( $\mu S$ )	Type of Bond
1	Sugar			
2	NaCl			
3	NaI			
4	KCl			
5	KI			
6	Distilled water			
7	Tap water			
8	Ethyl alcohol			

**Appendix 1.5: Bond Types and Conductivity** (*continued*)

---

- Predict the type of bond that each of the substances will be and record this in the appropriate column of the table.
- Place approximately 30 mL of sugar solution into a beaker. Label the beaker "sugar."
- Repeat step 3, using NaCl, NaI, KCl, and KI solutions.
- Prepare three beakers containing 30 mL of distilled water, tap water, and ethyl alcohol. Label them appropriately.
- Set up the computer system with the conductivity sensor.
- Calibrate the conductivity sensor.
- Display the conductivity tester with a digits display and/or meter display.
- Start the probeware data collection.
- Submerge the conductivity sensor end completely into the beaker of sugar solution.
- Once the reading has stabilized, record the measure of conductivity into the recording table.
- Remove the sensor from the beaker of sugar and completely rinse the sensor with distilled water. Then dry with tissue.
- Repeat steps 10 to 12 with each of the different liquids and solutions.
- Stop the data recording.
- Discard the solutions as directed by your teacher. Do not pour anything down the drain.

**Questions****Analysis**

- Which solutions were strong electrolytes and which solutions were weak electrolytes or non-electrolytes?
- Explain any differences in your results for distilled water and tap water.

**Conclusions**

- In general, what bond types conduct electricity?
- In general, what bond types do not conduct electricity?

Appendix 1.5: Bond Types and Conductivity (*continued*)

---

**Extensions**

5. How will the conductivity change if the amount of solute (one which has been shown to conduct electricity) dissolved into the solvent is increased to the point of saturation?
6. Solutions of ionic substance conduct electricity because a flow of electricity is caused by the charged ions in the solution. *Cations* (positive ions) move toward the negative electrode or cathode. *Anions* (negative ions) move toward the positive electrode or anode. This movement of ions toward the oppositely charged electrode in the solution causes the flow of electric charge. Why do cations move toward the cathode and not toward the anode?
7. Collect tap water from your home and test it for conductivity. How does it compare to the tap water from school or from another student's home?

## Appendix 1.5: Bond Types and Conductivity (Teacher Notes)

### Sample Data: Answer Key

#### Observations

Data Table				
Trial #	Liquid or Solution	Predicted Type of Bond	Conductivity ( $\mu\text{S}$ )	Type of Bond
1	Sugar		0.0	Covalent
2	NaCl		111.1	Ionic
3	NaI		112.8	Ionic
4	KCl		133.4	Ionic
5	KI		115.9	Ionic
6	Distilled water		0.0	Covalent
7	Tap water		42.3	Ionic
8	Ethyl alcohol		0.0	Covalent

#### Analysis

- Strong electrolytes: NaCl, NaI, KCl, and KI.

Weak electrolytes: tap water (depending on the water source—some water sources will be greater conductors of electricity than others).

Non-electrolytes: distilled water, sugar, and ethyl alcohol.

- Tap water varies in terms of number of free ions.  $\text{Fe}^{3+}$ ,  $\text{Ca}^{2+}$ ,  $\text{CO}_3^{2-}$ ,  $\text{Cl}^-$ , and traces of many others are often found in tap water. Distilled water usually has most of these ions removed.

#### Conclusions

- Ionic bonds conduct electricity.
- Covalent bonds do not conduct electricity.

#### Extensions

- Generally, the solution will conduct greater electricity as more solute is dissolved. At some point, the conduction will maximize and added dissolved solute will actually reduce the conductivity of the solution. Two different concentrations with the same conductivity measurement will result.
- Opposites attract and positive ions will be attracted by the negative electrode, while negative ions will be attracted by the positive electrode.
- Results will vary.

**Appendix 1.6: Chemistry Article Review Form**

Theme: \_\_\_\_\_

Find a newspaper/magazine/Internet article that fits a topic related to the chemistry theme above. Its minimum size must be 200 words.

Student Name: \_\_\_\_\_ Course: \_\_\_\_\_

Author(s) of Article: \_\_\_\_\_

Article Source/Website: \_\_\_\_\_

Article Title: " \_\_\_\_\_ "

Date of Article: \_\_\_\_\_ Page(s): \_\_\_\_\_

A. Attach the article (or photocopy) to the back of this page.

B. Highlight (or underline) the main idea found in the article. Briefly state the main idea in your own words: \_\_\_\_\_  
\_\_\_\_\_

C. List two important facts the author uses to support the main idea (full sentences):

1. \_\_\_\_\_

2. \_\_\_\_\_

D. Identify three science-related terms found in the article and write definitions (in context) for each:

1. \_\_\_\_\_ : \_\_\_\_\_  
\_\_\_\_\_2. \_\_\_\_\_ : \_\_\_\_\_  
\_\_\_\_\_3. \_\_\_\_\_ : \_\_\_\_\_  
\_\_\_\_\_E. Describe what you found most interesting about your article: \_\_\_\_\_  
\_\_\_\_\_  
\_\_\_\_\_

Remember, you are being assessed on all the above components of the review.

## Appendix 1.7: Making Fingerprints Visible

---

### Introduction

When we touch things, the ridges of our fingertips often leave impressions. These impressions are called fingerprints. We are not often aware that we leave fingerprints, and sometimes objects must be treated chemically to make these prints visible. In this activity, we will make a fingerprint on paper and then treat the paper to make the print visible. You might compare your fingerprints with those of other students. You will find that no two are alike.

### Materials

- unlined white index cards or white paper
- iodine ( $I_2$ ) crystals
- Erlenmeyer flask (250 or 125 mL)
- burner
- tongs
- scissors

### Procedure

1. Cut narrow strips of paper small enough to be held in the neck of the Erlenmeyer flask.
2. Press one of your fingers firmly on one end of a paper strip.
3. Place some iodine crystals (an amount about the size of a pencil eraser) in the flask. Heat gently with a burner or on a hot plate until the iodine begins to vaporize. **Caution:** Be careful. Do not breathe iodine vapour. Perform this in the fume hood.
4. Place the paper strip into the flask to expose it to the iodine vapour. Hold until changes to the paper are observed.

### Questions

1. The melting point of iodine is  $113.5^\circ\text{C}$ . What happens if iodine is heated above this temperature? (Ask the teacher for a demonstration.)
2. Compare the volume of iodine vapour to the volume of the same mass of solid iodine.
3. What changes of state took place in this experiment? Identify and describe each. *Example: "Freezing—the liquid water froze solid."*
4. What movie demonstrated a similar process? (Hint: Eddie Murphy)

## Appendix 1.8: Probeware Investigation: Determining Melting Points

### Question

How can the melting points of various organic compounds be found?

### Prediction

The melting point of a compound is the temperature at which the solid and liquid phases are in equilibrium. This is a physical property often used to identify compounds or to check the purity of the compound.

Predict what the melting points of different organic substances will be.

### Materials

- computer system and interface
- temperature sensor
- melting point capillary tubes (closed at one end)
- filter papers
- 1-gram samples from the following list: vanillin, naphthalene, resorcinol, acetanilide, benzoic acid, 2-naphthol, urea, maleic acid, cholesterol, citric acid, salicylic acid
- clean cloth
- Thiele tube filled with mineral oil to a level no more than 2 cm above the upper inlet of the sidearm
- two-holed stopper to fit the top of the Thiele tube
- 3 cm (3 mm outside diameter) rubber tubing
- scissors
- Bunsen burner

### Safety Precautions

- Be sure the room is well ventilated.
- This experiment must be done carefully to avoid burns and broken glass.

### Procedure

1. Set up the computer system with the temperature sensor set to record at a rate of once per second.
2. Represent the sensor with a graph (temperature vs. time) and digits display.
3. Obtain a melting point capillary tube and a sample of the known compound.
4. To load the melting point capillary tube, place a small amount of the compound on clean filter paper. Push the open end of the melting point capillary tube into the middle of the pile of compound. Some solid should be trapped in the tube—use the smallest amount of material that can be seen.
5. Turn the melting point capillary tube over, closed end down. Use a clean cloth to remove any compound sticking to the outside.

**Appendix 1.8: Probeware Investigation: Determining Melting Points** *(continued)*

---

6. Keeping the melting point capillary tube vertical, drop it from a height of 2 cm perpendicularly onto a solid surface repeatedly. The compound is packed when the entire compound has reached the bottom of the melting point capillary tube.
7. Clamp the Thiele tube just under the mouth with the utility clamp that is attached midway to the support stand.
8. Fit the temperature probe through the two-holed stopper, so that when placed into the Thiele tube, the tip will be immersed adjacent to the upper inlet of the sidearm.
9. Using scissors, cut a 2 mm section of the rubber tubing, to be used as a small rubber band.
10. Place the closed end of the loaded melting point capillary tube next to the end of the temperature probe. Place the rubber band around the temperature probe and melting point capillary tube so that it is 1 cm from the top of the melting point capillary tube.
11. Keeping the temperature probe vertical, place the stopper apparatus into the Thiele tube.
12. Start the temperature sensor to monitor the temperature.
13. Heat the mineral oil with a moderate burner flame, directing it at the curved side of the Thiele tube.
14. Allow the temperature to rise rapidly to within 15°C to 20°C below the expected melting point of the compound.
15. Adjust the flame size so the temperature rises no more than 2°C to 3°C per minute just before, during, and just after the period in which the compound melts.
16. Record the range of temperature from the first visible evidence of liquid (the sample appears moist, or a tiny drop of liquid is observed) to the complete liquefaction of the sample.
17. After the sample has melted, lift the thermometer and attached sample tube carefully (it may be hot) until they are just out of the oil. Wait for the thermometer to cool to about room temperature before you remove it entirely from the tube.
18. Remove the melting point capillary tube and wipe off some of the oil from the thermometer.
19. Reload a new melting point capillary tube (never re-melt melted samples), and repeat steps 3 to 18 for each of the rest of the samples provided.
20. Stop recording the data.
21. Clean up and discard the materials as directed by your teacher. Do not pour anything down the drain.

**Appendix 1.8: Probeware Investigation: Determining Melting Points** *(continued)*

---

**Questions****Analysis and Conclusions**

1. Why does the sample in the melting point capillary tube have to be packed tightly?
2. Why should the filled portion of the capillary tube be placed immediately adjacent to the mercury bulb of the thermometer?
3. What were the melting temperatures of the compounds analyzed?

**Applications**

4. List two ways in which the melting point of a solid organic compound could be useful to organic chemists.
5. What is the effect of a small amount of impurity on the melting point of an organic compound?
6. The freezing point of a substance has the same numerical value as its melting point, yet melting points are routinely measured but freezing points are not. Why?
7. Why is this method not used for finding the melting points of inorganic compounds?

## Appendix 1.8: Probeware Investigation: Determining Melting Points (Teacher Notes)

---

### Notes

1. If necessary, use a spatula to grind the compound into a fine powder so that it may enter the melting point capillary tube.  
Pack the compound into the melting point capillary tubes. If left loose, the compound will heat unevenly.  
A dropping tube may be used to assist in packing the compound into the melting point capillary tubes.
2. With the Thiele tube, hot oil is used to transfer heat evenly to your sample in the melting point capillary tube. Oil in the sidearm is heated and it expands to become less dense. The hot oil goes up the sidearm, warming the sample and the thermometer as it touches them. The oil cools and becomes denser and it falls to the bottom of the tube where it is heated again. This cycle goes on automatically as you do the melting point test in the Thiele tube.  
Water must not get into the Thiele tube. If it happens, the water can boil and throw hot oil out at you.
3. Take care that the rubber band securing the melting point capillary tube to the temperature probe remains above the mineral oil throughout the experiment.
4. Never re-melt any samples. They may undergo chemical changes such as oxidation, rearrangement, and decomposition.
5. The melting point is defined as the temperature range over which a small amount of solid in a thin walled capillary tube first visibly softens (first drop of liquid) and then completely liquefies. Thus, the melting point is actually a melting range.
6. Melting points recorded in chemical journals are capillary melting points unless otherwise stated.

### Answer Key

#### Analysis and Conclusions

1. So that the melting can be observed within in the same place where the temperature is being monitored.
2. This allows the melting temperature of the sample to be monitored with accuracy.

---

**Appendix 1.8: Probeware Investigation: Determining Melting Points  
(Teacher Notes) (continued)**

---

3. Answers will vary. Suggested results:

Compound	Melting Point (°C)
vanillin	77-82
naphthalene	79-80
resorcinol	109-110
acetanilide	113-114
benzoic acid	121-122
2-naphthol	121-122
urea	132-133
maleic acid	136-137
cholesterol	148-150
citric acid	150-153
salicylic acid	156-158

**Applications**

4. a) Determination of purity. Impurities generally have two effects: they lower the melting point from what it would be for the pure compound and they cause the melting point range to broaden.  
b) The melting point of an unknown sample can also be found by having students repeat the experiment and comparing their findings to a list of known melting points of different organic compounds. Many texts contain tables of melting points and lists of compounds that may have a particular melting point. One of them may be the unknown. If nothing else, you know the melting point of your unknown.
5. The melting point temperature range will be wider than 2°C.
6. In practice, freezing points are rarely measured because they are more difficult to determine than melting points. One reason for this is that solidification may not occur at the correct temperature due to the phenomenon of supercooling.
7. Inorganic compounds have extremely high melting points.

## Appendix 1.9: Vapour Pressure with Pop

---

### Introduction

A small volume of a liquid is added into a container and sealed. The container is then rolled to coat its sides with the liquid. After a minute or so, the cap pops off the container with report.

Chemical concepts: vapour pressure, volatility, Dalton's Law of partial pressures, and atmospheric pressure.

### Materials

- petroleum ether or pentane
- tennis-ball can (or a vial or a film canister)
- thin-stem polyethylene transfer pipette

### Safety Precautions

Petroleum ether and pentane form explosive mixtures with air. Do not perform this demonstration near a flame or near operating machinery. Aim the container into the air to avoid hitting anyone with the flying lid.

### Procedure

1. Add 3 to 4 mL of petroleum ether or pentane to a clean, dry tennis-ball can with a thin-stem pipette or by pouring a pre-measured amount from a beaker. Quickly seal the can with its lid and roll to coat the sides with the hydrocarbon. It is helpful if you warm the can with your hands as you roll it to increase the vapour pressure of the liquid. Have a student feel the sides of the can to observe the lower temperature, implying evaporation. After a minute or so, the lid will pop off the can with a loud noise.
2. This demonstration can also be performed using a vial or a film canister and pentane. Add enough liquid to cover the bottom of a large vial and quickly seal it by pressing on its lid. Swirl the container to coat the sides with the liquid, simultaneously warming it with your hands. If the lid does not pop in a minute or so, start the lid by carefully lifting it with your thumb. The experiment can often be repeated by simply pressing the lid back on the vial or adding more pentane.

### Notes

- The tennis-ball cans with the tightest fitting lids work best in this demonstration.
- The larger the vial is the better, since large vials have a proportionally small surface area in contact with the lid.
- Only those vials or film canisters with press-fit lids (the lids fit totally inside the vials) seem to work with any consistency.

---

**Vapour Pressure with Pop:** Copyright © by Walter Rohr, Eastchester High School, Eastchester, NY. Adapted by permission of the author. All rights reserved.

---

**Appendix 1.9: Vapour Pressure with Pop (continued)**

---

**Discussion**

Pentane and petroleum ether (a mixture of pentane, and various isomers of hexane) have vapour pressures in excess of 400 torr at room temperature (consult the Material Safety Data Sheet for the exact vapour pressure of your petroleum ether). Since the inside diameter of a tennis-ball can is roughly 7 cm, which gives an area of  $38 \text{ cm}^2$  ( $6 \text{ in.}^2$ ), the net force exerted would be 46 pounds ( $6 \text{ in.}^2 \times 14.7 \text{ lb./in.}^2 \times 400/760 \text{ atm}$ ), more than enough to blow off the lid. Pentane, with an even higher vapour pressure of 512 torr, would exert a net force of roughly 59 pounds under the same conditions ( $6 \text{ in.}^2 \times 14.7 \text{ lb./in.}^2 \times 512/760 \text{ atm}$ ).

The inside diameter of a film canister used in the demonstration is 3.0 cm. A 3.0 cm diameter cap has an area of  $7.0 \text{ cm}^2$  ( $1.1 \text{ in.}^2$ ) and, when using pentane, would experience a force of about 11 pounds ( $1.1 \text{ in.}^2 \times 14.7 \text{ lb./in.}^2 \times 512/760 \text{ atm}$ ). Again, this would be more than enough force to pop the lid.

This demonstration can also be used to illustrate Dalton's Law of partial pressures. What is the pressure inside the container before the cap flies off? Using Dalton's Law, the total pressure inside the container is the sum of the individual partial pressures.

$$\begin{aligned}\text{Pressure inside the container} &= P_{\text{atm}} + P_{\text{vapour}} = 760 \text{ torr} + 400 \text{ torr} \\ &\approx 1160 \text{ torr or } 1.5 \text{ atm}\end{aligned}$$

$$\begin{aligned}\text{Force exerted on the inside of the lid} &= 1.5 \text{ atm} \times 14.7 \text{ lb./in.}^2 \times 6 \text{ in.}^2 \\ &= 135 \text{ lbs.}\end{aligned}$$

What is the net force pushing the lid off the container?

$$\begin{aligned}\text{Net force} &= \text{Force inside the container} - \text{Force of the atmosphere} \\ &= 1.5 \text{ atm} - 1.0 \text{ atm} = 0.5 \text{ atm} \\ &= 0.5 \text{ atm} \times 14.7 \text{ lb./in.}^2 \times 6 \text{ in.}^2 = 46 \text{ lbs.}\end{aligned}$$

**Note:** The same calculations could be done with kPa but would not be as meaningful to most students.

**Disposal**

Pour the excess petroleum ether or pentane into a shallow pan or glass dish and allow it to evaporate in a working fume hood.

## Appendix 1.10: Measuring the Vapour Pressure of a Liquid (Student Experiment)

### Problem

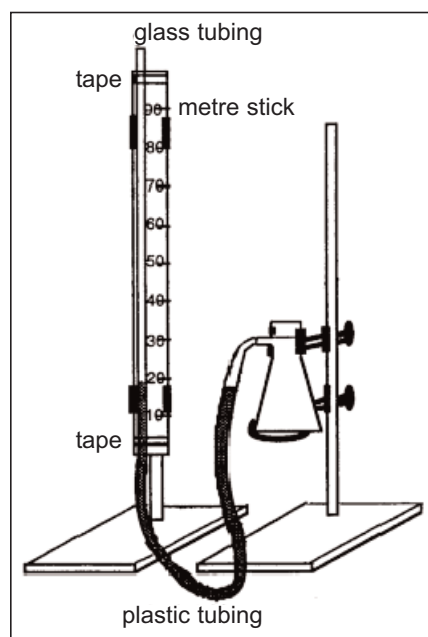
To measure the vapour pressure of a liquid.

### Materials

- samples of several liquids
- about 1 m flexible plastic tubing to fit glass tubing
- about 1 m glass tubing
- suction flask (250 mL)
- rubber stopper to fit suction flask
- 2 ring stands
- 1 ring clamp
- 3 clamps
- metre stick
- transparent tape
- graduated cylinder
- thermometer

### Procedure

1. Set up the apparatus as shown in the diagram to the right. **Note:** Fill the plastic tubing with water before attaching it to the suction flask and glass tubing.
2. Adjust the apparatus so that the level of water in the two arms is equal, and the level can be read on the metre stick. Read and record the water level.
3. Measure and record the temperature of the room.
4. Select one liquid. Add \_\_\_ mL of your liquid. Pour the liquid into the suction flask and quickly stopper the flask with the rubber stopper.
5. Describe what happens to the water in the tubing. Wait until no further change is apparent.
6. Adjust the apparatus so that the level of water in the two arms of the tubing can be read. Read and record the water level in *both* arms of the tubing.
7. What is the vapour pressure of the liquid? How do you know?
8. Either cool the outside of the suction flask by placing an ice cube on it, or warm the flask with your hands. Record your observations.



**Appendix 1.10: Measuring the Vapour Pressure of a Liquid  
(Student Experiment) *(continued)***

---

**Questions**

1. Is the vapour pressure of the liquid the same at all temperatures? How do you know? How could you find out? Can you devise a method to measure the vapour pressure of your liquid at 0°C?
2. Convert the vapour pressure you measured to vapour pressure in the following units: mmHg, torr, atmosphere, and pascal.
3. Why do you have to put a stopper on the flask?

## Appendix 1.10: Measuring the Vapour Pressure of a Liquid (Discussion—Teacher Notes)

The experiment described in Appendix 1.10 is simple and quick. For best results, give several students different amounts of the same liquid (e.g., 2 mL, 5 mL, 10 mL cyclohexane). Students should be able to deduce the following from the results:

- Time is required for vapour equilibrium to be established – the water level changes over two to five minutes. Students may not use the term “equilibrium” in their explanations; instead, they may say that some time is required for the apparatus to stabilize as the liquid vaporizes and mixes with the air in the flask.
- The vapour pressure is independent of the amount of liquid, provided some liquid is present in the container.
- Vapour pressure can be measured only in a closed system.
- Vapour pressure is different for different liquids.
- Vapour pressure varies with temperature.

Encourage students to set up their own apparatus. A #6 rubber stopper fits a 250 mL suction flask. Five-sixteenths (5/16th) inch Nalgene® or Tygon® tubing fits the sidearm of a 250-mL suction flask. Filling the tubing with water may be tricky – fill the tubing by slowly adding water and allowing the air bubbles to escape.

Some results obtained with this apparatus are:

Liquid	Temperature	“Height” in Glass	“Height” in Plastic	Vapour Pressure
acetone	22°C	62.2 cm	6.7 cm	55.5 cm water (40.8 mmHg)
ethanol (95%)	20°C	26.4 cm	4.8 cm	21.6 cm water (15.9 mmHg)
cyclohexane	20°C	39.5 cm	3.7 cm	35.8 cm water (26.3 mmHg)

### Notes:

- Both acetone and ethanol are highly soluble in water. Their vapours also dissolve, and the results tend to be low. Best results will be obtained with non-polar liquids such as cyclohexane or hexane.
- This experiment could easily be extended to measure the time to reach equilibrium, and so the rate of vaporization.
- Alternative methods may be used.
- An experiment to measure the vapour pressure of water as a function of temperature is also described in the following article:

Levinson, Gerald S. “A Simple Experiment for Determining Vapour Pressure and Enthalpy of Vaporization of Water.” *Journal of Chemical Education* 59.4 (April 1982): 337–338.

## Appendix 1.11: Forces between Particles

### Question

What is the relationship between the pressure and the temperature of a gas?

### Prediction

Once the vapour pressure curve of one liquid is determined, predict how it will compare to a different liquid's vapour pressure curve.

### Materials

- computer system and temperature pressure sensor
- two 250 mL pressure flasks (thick-walled)
- two-holed stopper (to fit into the flasks)
- glass eyedropper stem (glass part only)
- 800 mL beaker with hot water bath (80°C)
- hot plate
- 10 mL acetone,  $\text{CH}_3\text{COCH}_3$
- 10 mL ethyl alcohol,  $\text{C}_2\text{H}_5\text{OH}$

### Safety Precautions

- Wear a lab apron and safety goggles.
- Acetone is extremely flammable. Vapour is an irritant to eyes, skin, and lungs.
- Ethyl alcohol is highly flammable. High concentrations of vapour and concentrated solutions are dangerous.
- This experiment must be done carefully to avoid burns and broken glass.

### Procedure

1. Set up the computer system with the temperature sensor and pressure sensor.
2. Display the pressure sensor with a graph and the temperature sensor on the same graph so that pressure is recorded along the vertical axis and temperature is recorded along the horizontal axis.
3. Insert the thermometer sensor into one of the two holes of the stopper (from the wide end out through the narrow end of the stopper) so that its tip will be within the centre of either flask when it is stoppered.
4. Insert the wide end of the glass eyedropper stem into the other outer hole of the wide end of the stopper.
5. Connect the pressure sensor tube to the narrow tip of the glass eyedropper.
6. Prepare the beaker with the hot water bath  $\frac{3}{4}$  full at 80°C.
7. Place 10 mL of acetone into a 250 mL flask.
8. Place the flask into the hot water bath for five minutes.
9. Begin monitoring the temperature and pressure.

**Appendix 1.11: Forces between Particles** (*continued*)

---

10. Insert the two-holed stopper with pressure and temperature sensor into the flask.
11. Remove the flask from the hot water bath.
12. Once the temperature and pressure stabilize, stop monitoring the data and start recording the data for Run #1.
13. Stop the data recording once the graph reaches 30°C.
14. Clean up the materials as directed by the teacher.
15. Repeat steps 7 to 14, using a clean flask and ethyl alcohol instead of acetone.

**Questions****Analysis**

1. What experimental factors were kept constant during this experiment?
2. How are the graphs of pressure versus temperature the same between acetone and ethyl alcohol?
3. How are the graphs of pressure versus temperature different between acetone and ethyl alcohol?

**Conclusion**

4. Based on the data and graph that you obtained for this experiment, what is the general relationship between pressure and temperature?
5. Is this relationship direct or indirect?

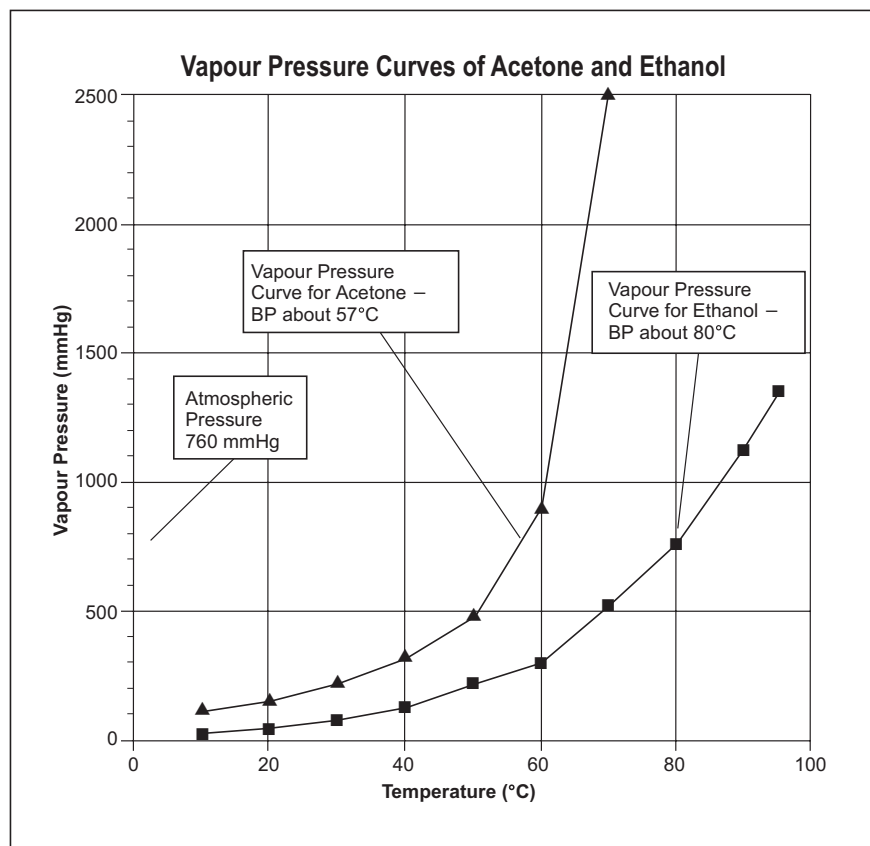
**Applications**

6. Based on your vapour pressure measurements at room temperature, which compound do you predict has stronger intermolecular forces at work? Explain your reasoning.
7. Define the term *volatile*. Which of the two liquids do you expect to be the least volatile? Why?
8. Why are most pressurized cans labelled with warnings not to dispose of the cans by throwing them into a fire?

## Appendix 1.11: Forces between Particles (Teacher Notes)

### Sample Data: Answer Key

#### Observations



#### Analysis

1. The experimental factors that were kept constant during this experiment include the volume of the flask and the purity of the acetone and ethyl alcohol.
2. Both graphs have a similar sloping shape.
3. Acetone's graphed curve has a steeper slope.

#### Conclusion

4. As the temperature of a gas is decreased, its pressure also decreases. As the temperature of a gas is increased, its pressure also increases.
5. This relationship is direct.

Appendix 1.11: Forces between Particles (Teacher Notes) *(continued)*

---

**Applications**

6. Ethyl alcohol has stronger intermolecular forces at work. A larger amount of energy is required to overcome the stronger intermolecular forces that hold the liquid together.
7. *Volatile* means that a substance is easily vaporized at a lower temperature. Acetone is more volatile than ethyl alcohol.
8. With extremely high temperatures created by fire, the contents of pressurized cans will also become very high, resulting in possible explosions.

## Appendix 1.12: Freezing by Boiling

### Introduction

When a small amount of a liquid is added to a flask and evacuated at reduced pressure, the liquid boils. After a minute or so, the liquid freezes while it is boiling.

### Chemical Concepts

- vapour pressure
- boiling point
- triple point
- heat of vaporization

### Materials

- cyclohexane
- Erlenmeyer flask (125 mL) or a large test tube
- one-hole stopper to fit the flask or test tube
- medicine dropper
- aspirator with a splash guard
- 3 to 4 ft. of 3/16 in. (internal diameter) vacuum tubing
- boiling chips

### Procedure

1. Add 30 to 40 mL of cyclohexane and a boiling chip to a 125 mL Erlenmeyer flask or large test tube. Seal the flask or test tube with a rubber stopper attached to an aspirator. The glass part of a medicine dropper makes a convenient connector between the rubber stopper and the vacuum tubing.
2. Turn the water on completely and observe. Challenge a student to touch the test tube while it is boiling. Students believe the test tube will become hot. Continue to evacuate the system until the cyclohexane is completely frozen and no additional change occurs. Pass the flask containing the frozen cyclohexane around the room so that students can observe the temperature of the flask. The same sample of cyclohexane can be boiled again and the demonstration repeated if another boiling chip is added to the flask.

### Discussion

The aspirator lowers the pressure exerted on the cyclohexane until the vapour pressure of the liquid equals the applied pressure where the liquid boils. Since boiling is an endothermic process, the heat of vaporization is absorbed from its own kinetic energy, causing the temperature of the cyclohexane to drop until its freezing point of  $6.6^{\circ}\text{C}$  is reached. The cyclohexane will continue to boil while it is freezing at its triple point.

**Appendix 1.13: Gas Laws: Temperature and Pressure Changes**

---

**Question**

How does a decrease in pressure above a liquid affect its boiling point?

**Prediction**

Predict whether lowering the pressure of the atmosphere above a liquid will raise or lower its boiling point.

**Materials**

- computer system and temperature and pressure sensor
- 250 mL boiling flask (thick-walled)
- two-hole stopper (to fit into the flask)
- glass eyedropper stem (glass part only)
- distilled water
- boiling chips
- hot plate
- gloves or an insulating cloth
- support stand
- utility clamp
- a large pail of cold water
- utility pan (large enough to contain the pail's contents of water)

**Safety Precautions**

- Wear a lab apron and safety goggles.
- This experiment must be done carefully to avoid burns and broken glass.

**Procedure**

1. Set up the computer system with the temperature sensor and pressure sensor.
2. Display the temperature sensor with a graph (temperature vs. time) and the pressure sensor with a separate graph (pressure vs. time).
3. Add 10 boiling chips to the boiling flask and fill it one-third full of distilled water.
4. Insert the thermometer sensor into one of the two holes of the stopper so that its tip will be within the water in the flask when it is stoppered.
5. Insert the wide end of the glass eyedropper stem into the other outer hole of the stopper.
6. Connect the pressure sensor tube to the narrow tip of the glass eyedropper.
7. Set the support stand inside the utility pan and attach the utility clamp 30 cm from the base.

---

**Appendix 1.13: Gas Laws: Temperature and Pressure Changes** (*continued*)

---

8. Heat the boiling flask (without the stopper) on the hot plate until the water is boiling.
9. Start recording the data.
10. Remove the distilling flask from the heat and immediately insert the stopper into the flask. The stopper must be inserted just before the water stops boiling. If the stopper is put in too soon, the pressure of the steam will blow it out. If it is put in too long after the boiling has stopped, air will have entered the flask and the experiment must be restarted.
11. Attach the flask to the utility clamp.
12. Slowly run cold water from the pail over the flask and observe any boiling.
13. Continue until there is no more water to pour over the flask or the water in the flask no longer boils.
14. Stop recording the data.
15. Clean up the materials as directed by your teacher.

**Questions****Analysis**

1. What was present above the water in the flask just before the stopper was inserted?
2. What happened to the substance when cold water was run over the flask?
3. How did the cold water poured over the flask create a lowered pressure inside of the flask?
4. After the first of the cold water was poured, what was the temperature of the water inside the flask?
5. At what temperature did you stop pouring water or at what point did the boiling stop?
6. What can you conclude about the effect of lowering pressure on boiling?

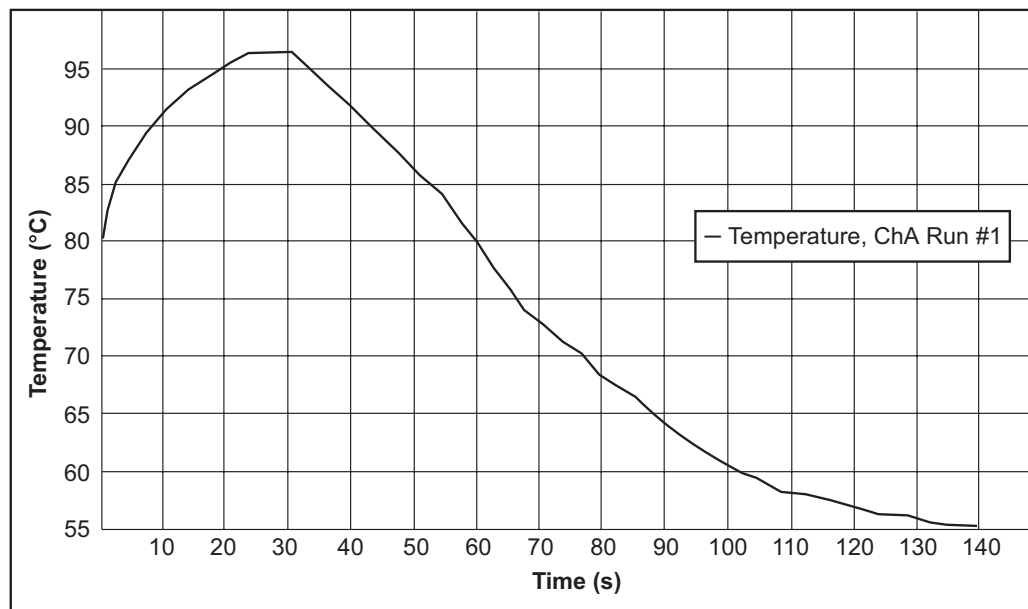
**Applications**

7. What is the name of the gas law that applies to both temperature and pressure? What is the formula?
8. Is the relationship between pressure and temperature direct or indirect?
9. At higher and higher altitudes the pressure decreases. Why would cooking potatoes in open pots not work at high altitudes?
10. Suggest a practical method of boiling and cooking potatoes at high altitudes.

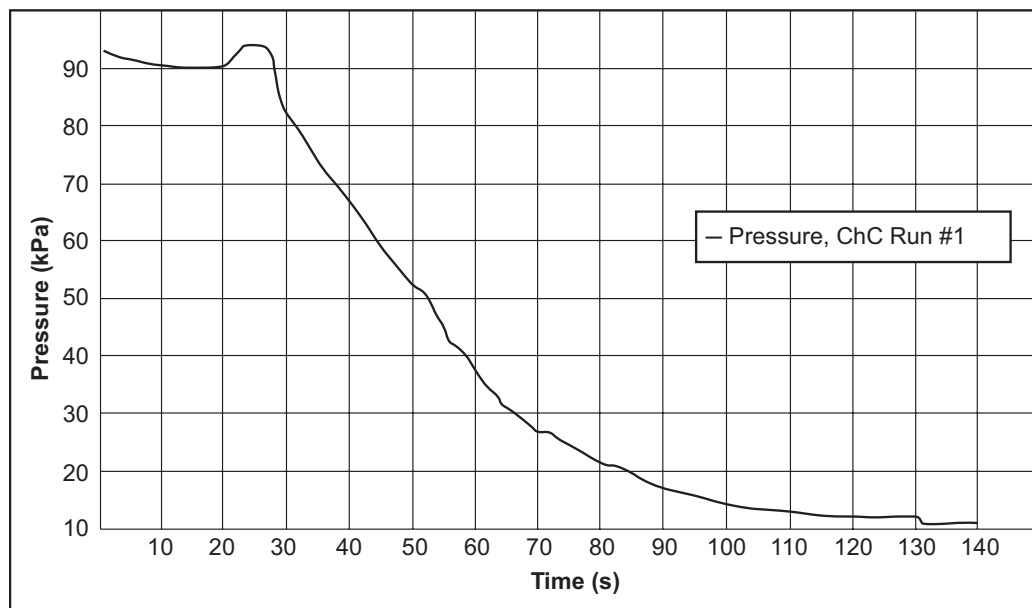
## Appendix 1.13: Gas Laws: Temperature and Pressure Changes (Teacher Notes)

### Sample Data: Answer Key

#### Observations

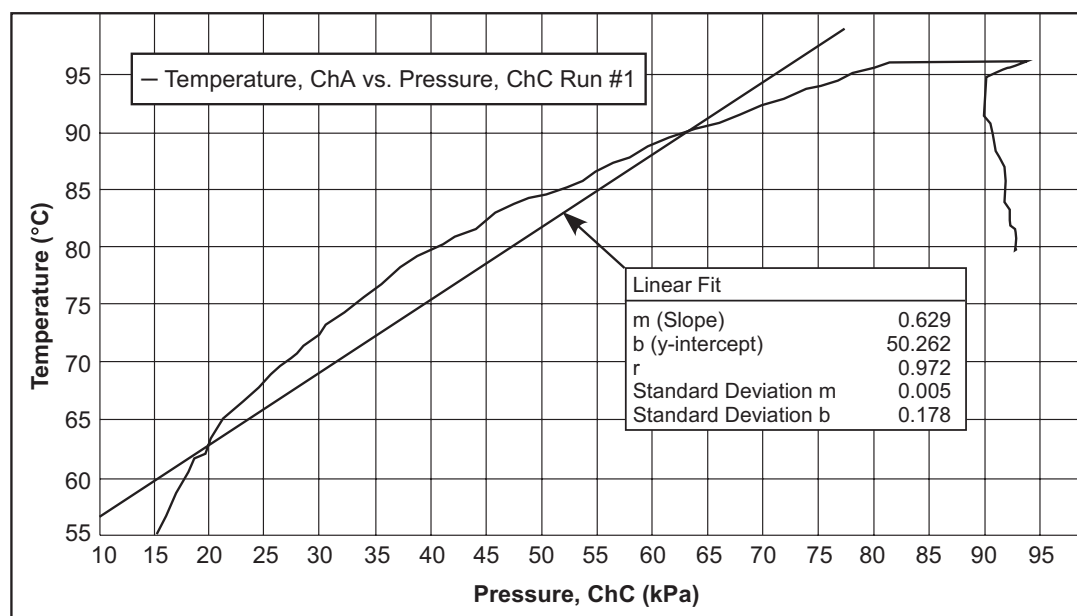


Graph 1: Temperature vs. Time



Graph 2: Pressure vs. Time

### Appendix 1.13: Gas Laws: Temperature and Pressure Changes (Teacher Notes) (continued)



Graph 3: Temperature vs. Pressure

#### Analysis

1. Steam.
2. The substance condensed onto the inside of the flask.
3. Condensation requires the vapour particles to condense, resulting in a partial vacuum.
4. Around 99°C.
5. Around 60°C (depending on the temperature and amount of water in the pail).
6. Lowering pressure lowers the boiling point of a liquid.

#### Application

7. Gay-Lussac's Law. 
$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$
8. The relationship between pressure and temperature is direct.
9. The boiling temperature would not be hot enough to cook the potatoes.
10. Through the use of a pressure cooker.

**Appendix 1.14: Chemistry is Super: “Bingo” Review Game**

---

**Time:** 1 hour

**Purpose**

To review the learning outcomes and concepts addressed in Topic 1: Physical Properties of Matter.

**Game Instructions**

1. Supply each student with a minimum of two slips of paper.
2. Assign each student two learning outcomes/concepts from Topic 1. (Vary the outcomes from student to student so that all outcomes/concepts for this topic are addressed.)
3. Ask students to create a question related to each of their assigned learning outcomes/concepts, and explain that the questions will be used in the review game to follow. Each question must be written in such a way that the answer will be a one-word response (i.e., fill-in-the-blank). The answer should be included on the slip of paper, along with the question. (In order for the game to work, at least 40 questions and answers must be provided.)
4. As students finish their questions, collect and verify them, and start to create a Word Splash of all the answers on the classroom board. (Time allotment up to this point should be no more than 30 minutes.)
5. When all slips have been submitted, and an entire Word Splash is on the board, hand out the Chemistry Is Super game sheet.
6. Have students randomly place 25 of the words written on the board onto their game sheet (one word per square). If two words are the same, students can put it down twice on their game sheet if they want to.
7. The teacher (or student) can read the questions from the slips one at a time to the class. As students determine the correct answer to a question, they cross out the answer on their game sheet.
8. When a student successfully crosses out a line (any direction), he or she must clearly and enthusiastically call out “Chemistry Is Super” to win. Play can continue until the end of class and multiple winners can be awarded, if desired.
9. Answers can be shared or discussed within the class for additional review. The game can also be played silently, with individual students trying to determine the answers on their own.

**Game Sheet**

A copy of the Chemistry Is Super game sheet is provided on the following page.

Appendix 1.14: Chemistry is Super: "Bingo" Review Game (*continued*)

---

<b>Chemistry Is</b>				
<b>S</b>	<b>U</b>	<b>P</b>	<b>E</b>	<b>R</b>

