# How Cool Is That?

## Introduction

The "cooling effect of evaporation" is nature's most important way of cooling our bodies and also the Earth. How cool is that? The purpose of this technology activity is to measure temperature changes versus time when different liquids evaporate. Liquids will be compared pair-wise (e.g., polar versus nonpolar) to analyze the strength of attractive forces.

## Concepts

- Evaporation
- Kinetic-molecular theory
- Intermolecular forces

## Background

Vaporization or evaporation is the process by which a substance changes from a liquid to a gas or vapor. Evaporation is an endothermic process—energy is required for molecules to leave the liquid phase and enter the gas phase. When the heat energy for vaporization comes from the surroundings rather than from external heating, the temperature of the surroundings will decrease when a liquid evaporates. This is the origin of the cooling effect of evaporation and explains why water evaporating from the skin by perspiration cools the body.

Evaporation and the cooling effect of evaporation may be explained using the *kinetic-molecular theory*. The temperature of a substance is proportional to the *average kinetic energy*, and thus the average speed, of the molecules. Evaporation occurs when fast-moving molecules near the surface of a liquid have enough energy to break free of their interactions with neighboring molecules and "escape" into the gas phase. Molecules with the highest kinetic energy evaporate and become gas molecules. The average kinetic energy, and thus the temperature, of the remaining molecules decreases—a liquid cools as it evaporates. This phenomenon is known as "evaporative cooling." The rate of evaporation of a liquid increases at higher temperatures, because more molecules have enough energy to break free of the liquid's surface.

The rate of evaporation of a liquid depends on the nature of the liquid and the type of attractive forces between molecules. Strong intermolecular attractions hold the molecules in a liquid more tightly. Liquids with weak intermolecular attractive forces are volatile—they evaporate easily. Liquids with strong intermolecular attractive forces are less volatile because a greater amount of energy is needed to overcome the attractive forces between the molecules.

Nonpolar compounds generally have weak attractive forces, called London dispersion forces, between molecules. The strength of London dispersion forces increases as the size of the molecules increase. Dipole interactions occur when polar molecules are attracted to one another and are usually stronger than dispersion forces. Hydrogen bonding represents a special case of dipole interactions, in which F–H, O–H and N–H groups in molecules associate with electronegative atoms in adjacent molecules. Hydrogen bonds are generally the strongest type of intermolecular attractive force.

## Materials *(for each student group)*

Acetone, $(CH_3)_2CO$, 2 mL
Heptane, $C_7H_{16}$, 2 mL
Hexane, $C_6H_{14}$, 2 mL
Isopropyl alcohol, $(CH_3)_2CHOH$, 2 mL
Computer interface system and software
Corks or stoppers to fit test tubes, 4
Filter paper or cotton gauze, 11-cm
Pipets, 4
Rubber bands, small (orthodontic-type), 4
Scissors
Temperature probes or sensors, 2
Test tubes, small, 4
Test tube rack

## *Safety Precautions*

*The solvents used in this activity are flammable liquids and dangerous fire risks. Avoid contact of all liquids with heat, flames or other sparks. Hexane is a suspected reproductive hazard. Hexane and heptane may cause drowsiness or dizziness if inhaled. Perform this experiment in a hood or well-ventilated lab only. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory. Please review current Safety Data Sheets for additional safety, handling, and disposal information.*


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## Procedure

1. Cut four small pieces of filter paper or cotton gauze, approximately 3 cm square each.
2. Label four test tubes #1–4 and place them in a test tube rack. Obtain 2 mL of the appropriate solvent in each tube and stopper with corks or rubber stoppers until needed.

| **Test Tube** | **1** | **2** | **3** | **4** |
|---|---|---|---|---|
| **Solvent** | Hexane | Heptane | Acetone | Isopropyl alcohol |

3. Wrap one filter paper square around each temperature probe and secure the filter paper with a small rubber band.
4. Plug the temperature probes (sensors) into an appropriate interface system and connect to a computer or calculator as needed.
5. Place one temperature probe into test tube #1 (hexane) and the other into test tube #2 (heptane). The liquid level in each test tube should be above the filter paper to thoroughly soak the paper. Allow the temperature probes to soak in the liquid for about 30 seconds.
6. Begin collecting data (temperature versus time) at 2 second intervals with the temperature sensors. Remove the sensors from the liquid after 10 seconds and extend the sensors over the test tube rack as shown in Figure 1. Continue collecting data for 180 seconds as the liquids evaporate from the probes. *(This is Trial A.)*
7. Remove the filter paper squares from the temperature probes and carefully dry the probes with a paper towel.
8. Repeat steps 3–7 using acetone and isopropyl alcohol. *(This is Trial B.)*

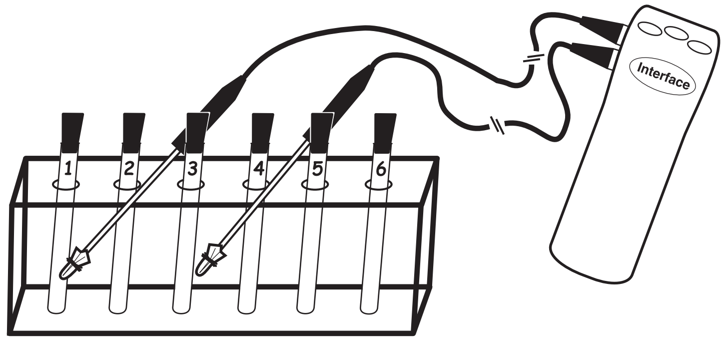



**Figure 1.**

## Questions

1. Describe a typical temperature versus time graph for the evaporation of a liquid and explain in terms of the cooling effect of evaporation.
2. Compare the temperature changes for hexane and heptane in Trial A. Which compound evaporated more quickly? How are these compounds similar? How are they different? Which compound has stronger attractive forces? Explain.
3. Compare the results obtained for acetone and isopropyl alcohol in Trial B. Which compound evaporated more quickly? How are these compounds similar? How are they different? Which compound has stronger attractive forces? Explain.
4. Rank the four liquids tested from most volatile to least volatile based on the observed temperature changes.

## Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. There should be no residual solvents.

## Connecting to the National Standards

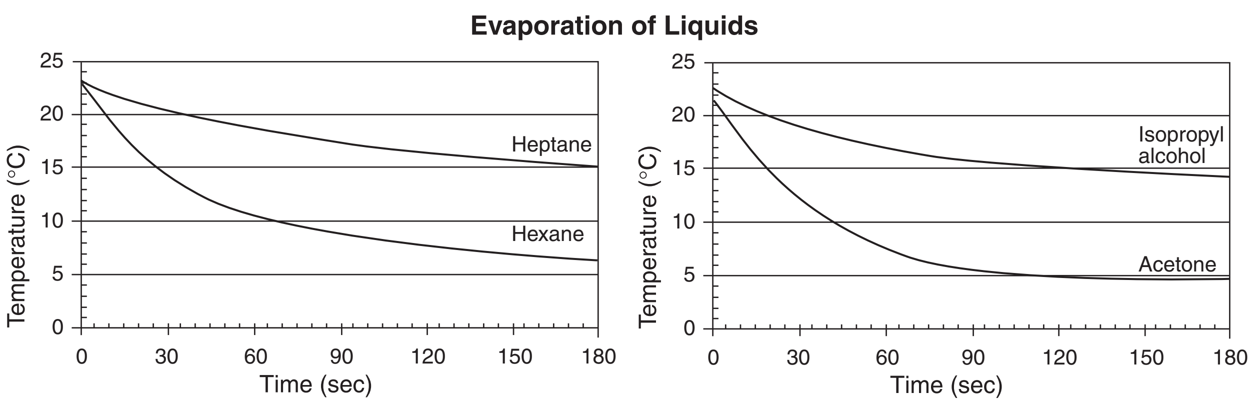
This laboratory activity relates to the following National Science Education Standards (1996):

***Unifying Concepts and Processes: Grades K–12***
Systems, order, and organization
Evidence, models, and explanation
Constancy, change, and measurement
***Content Standards: Grades 9–12***
Content Standard A: Science as Inquiry
Content Standard B: Physical Science, structure and properties of matter

## Tips

- Each trial takes only a few minutes to set up and collect data. Adjust the working group size as needed to accommodate the number of technology stations or temperature probes available in the lab.
- "Hexanes," a mixture of *n*-hexane and other isomers, may be used as the source of hexane for Part A. The boiling point of the mixture (68–70 °C) is very close to that of *n*-hexane (69 °C). In general, branched-chain alkanes have lower boiling points than their straight-chain isomers—the boiling points of *n*-hexane, 3-methylpentane (1 branch) and 2,3-dimethylbutane (2 branches) are 69 °C, 63 °C, and 58 °C, respectively.
- Isopropyl alcohol is recommended rather than *n*-propyl alcohol (1-propanol) because it is more readily available and is less toxic by inhalation (the TLV is 983 mg/m$^3$ for isopropyl alcohol and 492 mg/m$^3$ for *n*-propyl alcohol).
- In comparing the strength of attractive forces between different types of molecules, it is important to keep the size of the molecules or their molar mass similar. Dispersion forces increase for all types of molecules as they increase in size.
- This experiment may also be done using digital thermometers and a watch or timer with a second hand. Measure the temperature at least every 10–15 seconds.
- Use glass Pasteur pipets or medicine droppers for transferring organic solvents.
- Many different solvent pairs may be tested in this experiment—please consult SDS for the inhalation hazards. Pentane is not recommended because it has a very low flash point and is narcotic in high concentrations.
- Evaporative cooling is an energy-efficient method of cooling homes in hot and dry climates. A typical evaporative cooler uses about one-fourth as much electricity as an air conditioner and is more environmentally friendly. A large fan draws outside air through a water-soaked pad—the resulting evaporation of water cools the air and increases its moisture content. The resulting temperature decrease depends on the temperature and relative humidity of the incoming air.

## Sample Results



## Answers to Questions *(Student answers will vary.)*

1. Describe a typical temperature versus time graph for the evaporation of a liquid and explain in terms of the cooling effect of evaporation.

*The initial temperature of the liquid is constant at or near room temperature for the first 10 seconds or so, corresponding to the first 3–4 data points, when the temperature probe is still in the liquid (see step 5 in the Procedure). There is then a sharp decrease in the temperature over the next 30 seconds as the liquid begins to evaporate. The temperature will usually continue to decrease over the entire time period (180 sec), although the rate at which the temperature decreases begins to slow down or level off after about 100 seconds. The minimum temperature is generally observed after about 150 seconds. The temperature decrease occurs because a liquid cools as it evaporates and absorbs heat from the probe. The rate of evaporation of a liquid decreases at lower temperatures. The rate at which the temperature decreases thus reflects the rate of evaporation—as the liquid cools, it evaporates more slowly.*

2. Compare the temperature changes for hexane and heptane in Trial A. Which compound evaporated more quickly? How are these compounds similar? How are they different? Which compound has stronger attractive forces? Explain.

*Hexane reached a lower minimum temperature than heptane. This means that hexane evaporated more quickly than heptane. Hexane and heptane have very similar structures—they are both nonpolar hydrocarbons consisting of C—C and C—H bonds. Heptane has one more –$CH_2$– group in its "chain" and thus has stronger attractive forces than hexane. (Dispersion forces are stronger for "bigger" molecules.)*

3. Compare the results obtained for acetone and isopropyl alcohol in Trial B. Which compound evaporated more quickly? How are these compounds similar? How are they different? Which compound has stronger attractive forces? Explain.

*Acetone reached a lower minimum temperature than isopropyl alcohol. This means that acetone evaporated more quickly than isopropyl alcohol. Acetone and isopropyl alcohol have similar molar masses and both are polar compounds. Isopropyl alcohol, however, has an –OH group in its structure and is thus capable of forming hydrogen bonds with neighboring molecules. Isopropyl alcohol has stronger attractive forces than acetone.*

4. Rank the four liquids tested from most volatile to least volatile based on the observed temperature changes.

*From most to least volatile: Acetone > hexane > isopropyl alcohol > heptane*

## Reference

This activity was adapted from an experiment in *Solids and Liquids*, Volume 11 in the *Flinn ChemTopic*™ *Labs* series, Cesa, I., Editor; Flinn Scientific, Inc., Batavia IL (2005).

## Materials for *How Cool Is That?* are available from Flinn Scientific, Inc.

| Catalog No. | Description |
|---|---|
| A0009 | Acetone, Reagent, 500 mL |
| H0051 | n-Heptane, 100 mL |
| H0046 | Hexanes, Reagent, 100 mL |
| I0019 | Isopropyl Alcohol, 500 mL |

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.