

Name _____ School _____ Date _____

Lab 12.1 – Calorimetry

Purpose

To investigate the flow of heat between two bodies due to a difference in their temperatures.
 To investigate the flow of heat involved in changes in phase.

Equipment

Virtual Calorimetry Lab
 PENCIL

Explore the Apparatus

Open the Virtual Calorimetry Lab.

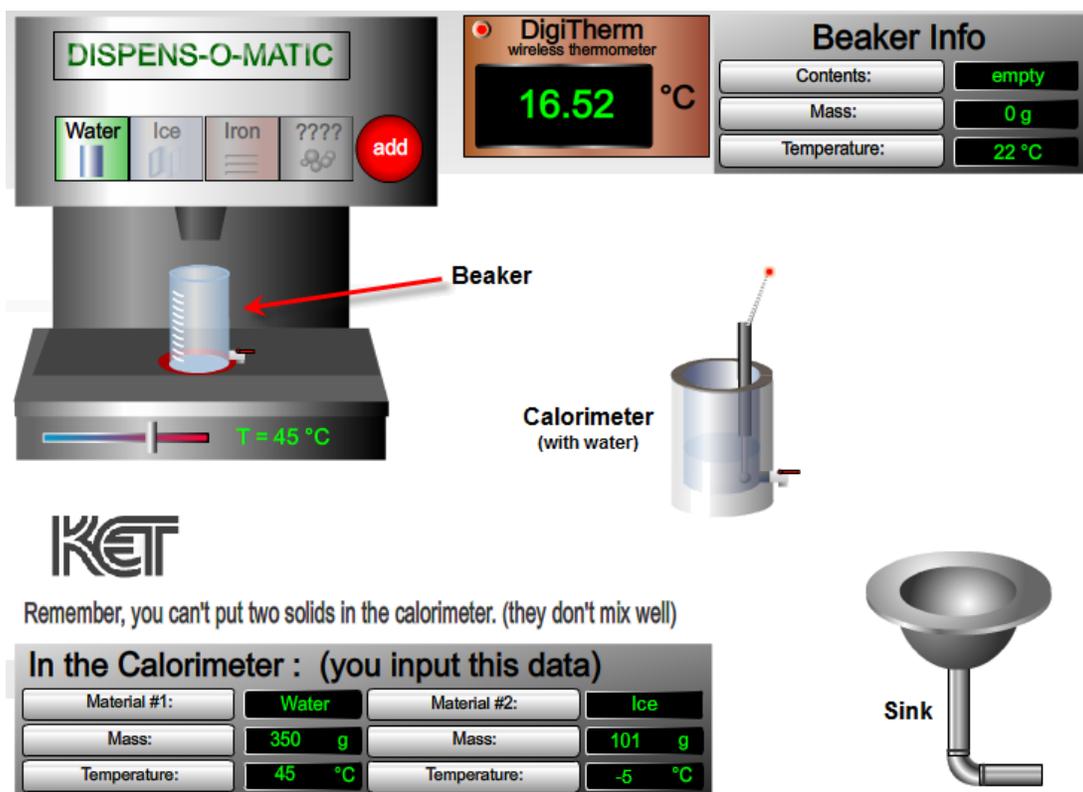


Figure 1

OK, it looks like we're not in mechanics anymore! Let's take a trial run with the calorimetry apparatus.

What we have is your basic Dispens-O-Matic which can dispense water, ice, iron, and a mystery material in a range of masses and temperatures. All the dispensing is done into the small graduated beaker. The beaker can receive only one type of material at a time. From there it can be transferred into the calorimeter or the sink for disposal.

Let's try to reproduce the situation depicted in Figure 1. Drag the beaker onto the hot/cold plate under the dispenser spout as shown in the figure. It will snap into place, even if you miss by a lot. Click the Water button. Drag the temperature slider to adjust the water temperature to 45 °C. Click and hold the Add button until the beaker is about half full.

Before continuing, you need to record your results so far in the "In the Calorimeter" table provided on the screen. Click in the text area beside Material #1. Type "Water." Continue by entering your mass and temperature as shown in the Beaker Info panel. You probably have a different mass than in the figure. The mass is hard to set precisely.

Let's transfer our water sample into the calorimeter now. Drag the beaker until your pointer is somewhere over the calorimeter. By "over" we mean on top of it on the screen. (Figure 2.)

When you release the mouse button the beaker will empty itself into the calorimeter.

Notice that the DigiTherm reading now gradually changes until it reads about 45 °C reflecting the temperature of the water in the calorimeter. The DigiTherm reading always reflects the temperature in the calorimeter. The Beaker Info box has now changed to "empty," "0 g," and "22 °C" reflecting the room-temperature air now in the beaker. That's why you need to record your data before transferring the sample.

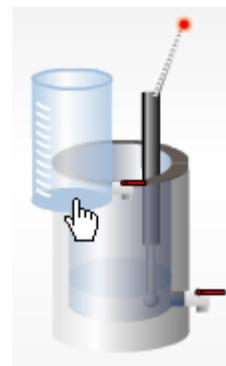


Figure 2

For material #2 we want about 100 g of -5 °C ice. Go ahead and dispense it. Record your data in the calorimeter data table and then add the ice to the calorimeter as before. You'll see all the ice gradually melt as the temperature gradually settles on an equilibrium temperature of about 16.5 °C.

You now have a complete set of data for the experiment. You have all the materials, masses, and temperatures before mixing as well as the final equilibrium temperature. You can see that all the ice is gone or infer it from the >0 °C temperature. The mass of the water in the calorimeter is just the sum of the two initial masses.

To clean up and get ready for the next trial, dump the calorimeter contents into the sink. Just drag the calorimeter until your pointer is anywhere over the sink. Incidentally, you can also dump materials straight from the beaker to the drain. You'll want to do that when you make a mistake such as when you add too much of something to the beaker. And make sure that you always empty your calorimeter before each new investigation.

Also note that you can't mix two solids in the calorimeter, so there will always be water involved. And you can just add materials to the calorimeter twice in a given trial. So you can't add ice, then water, and then add some more ice.

Initial Exploration

From your everyday experience you should have a pretty good idea about what happens when materials of different temperatures are mixed. Let's explore this a bit.

1. Suppose you mixed together approximately 100 g of 24 °C water with 100 g of 12 °C water. (The masses will always be approximate.)

What would you predict for the equilibrium temperature of this mixture? _____ °C

2. Try it with your calorimeter and see what actually happens. _____ °C

Hopefully you got 18 °C. If not, please try again.

3. No pressure here, but try to explain what might be happening that would explain this result.

4. Now suppose you mixed together approximately 200 g of 24 °C water with 100 g of 12 °C water.

What would you predict for the equilibrium temperature of this mixture? _____ °C

5. Try it with your calorimeter and see what actually happens. _____ °C

6. How's your theory holding up? You have two types of information – **mass** and **change in temperature**. See if you can come up with an equation using just those quantities that predicts the final temperature, T_f , of both of these trials. Show your calculations for each trial.

Theory

You've noticed that when samples of water at different temperatures are mixed, the resulting mixture has a temperature somewhere between the hot and cold temperatures of the original samples. Some sort of averaging seems to be taking place which would suggest the sharing of something. Investigating just how this happens on a microscopic scale is not the subject of this lab. But summarizing what is known about this behavior will provide us with both models and language to help us interpret what's happening.

Temperature vs. Heat

Try this. Find a desk or chair that has both metal parts and wooden or plastic parts. Hold the metal part with one hand while holding the wooden part with the other hand.

1. Which of these two materials appears to be warmer? metal wood

You found earlier that when two samples of water were mixed, they came to a common temperature. Both parts of the chair were probably in the same uniform-temperature room for some time, so wouldn't our mixing model suggest that they'd be the same temperature as the air?

Yes. And surprisingly they are the same temperature. So what gives? Consider this first.

2. Of the three objects involved, which is probably hotter? metal wood your hand

We go to a lot of effort to keep our bodies in a relatively cool environment. Our metabolism produces a lot of waste energy and we need to get rid of it somehow, thus our choice of a cool environment. When you touched the two different materials, they were initially probably about 15°C cooler than your body and as a result, heat flowed from you into both bodies. But the metal is a much better conductor of that heat, so when you sensed a lower temperature, what you were really detecting was a greater rate of flow of heat. The rate of flow of this heat is greater for the hand in contact with the metal.

In the preceding discussion we've used the term heat several times. Here's how we define heat.

Heat is the energy that flows from a high-temperature object to a low-temperature object because of a difference in temperatures. Heat will continue to flow until the materials reach **thermal equilibrium**. Or stated differently, two bodies are in a state of thermal equilibrium if no heat is exchanged between them when brought into thermal contact. At thermal equilibrium they will be at a common temperature, the **equilibrium temperature**.

The energy involved in this transfer is internal energy. The **internal energy** of a substance is the sum of the molecular kinetic energy, the molecular potential energy, and other kinds of molecular energy.

Defining temperature should help to pull these ideas together to explain what we've been observing with our water mixtures. While temperature is defined differently in different situations, for the present investigation we'll use the following. The **absolute temperature** of a body is a measure of the average internal energy per molecule.

Two bodies at the same temperature might have the same or different amounts of internal energy depending on the number of molecules. If one body contained more molecules, its total internal energy would be greater since the temperature represents the average internal energy per molecule.

Based on the above definitions we'd say that when we mix water of different temperatures together, internal energy is transferred by the interaction of the molecules. This energy in transfer from a warmer body to a cooler body is what we refer to as heat.

Specific Heat Capacity

When we mixed equal masses of 24 °C and 12 °C water the final temperature was the average of the two initial temperatures. If we had used water hotter than 24 °C, we would expect a higher final temperature since the higher initial temperature would represent a larger quantity of internal energy.

Similarly we found that when we doubled the amount of the warm water, we increased the final temperature of the mixture. Again this fits with our understanding of internal energy being related to the energies of the molecules. More molecules at a given temperature would correspond to more total internal energy.

So the more mass we have and the higher the temperature, the more internal energy there is to pass on to cooler materials as heat. Let's continue our initial exploration with one more investigation.

3. Suppose you mixed together approximately 100 g of 24 °C water with 100 g of 12 °C iron.

What would you predict for the equilibrium temperature of this mixture? Less than 18 °C 18 °C Greater than 18 °C

4. Try it with your calorimeter and see what actually happens. $T_f = \underline{\hspace{2cm}}$ °C

That's very different! Clearly water and iron are somehow not evenly matched. Here's a summary of what we've found.

| Table 1 Mixing Hot Water with Cold Water, and then Cold Iron | | | | | |
|--|------------|-----------|--------------------------|------------------------|----------------------------|
| Trial | | Mass (kg) | Initial Temperature (°C) | Final Temperature (°C) | Change in Temperature (C°) |
| 1 | Hot Water | .100 | 24 | 18 | -6 |
| | Cold Water | .100 | 12 | 18 | +6 |
| 3 | Hot Water | .100 | 24 | 18 | -1.2 |
| | Cold Iron | .100 | 12 | 22.8 | +10.8 |

The hot water in trial 3 cooled by only a small amount relative to the amount of cooling in trial 1. And the temperature of the cool metal changed by a much larger amount than the equal mass of cool water. From this we can surmise that a much smaller amount of heat is required to produce a given change the temperature of a given mass of iron.

In summary we've found that when heat flows into or out of a body there are (at least) three factors that determine the **amount of temperature change** – the **mass**, the **type of material**, and the **quantity of heat flowing**. We describe this mathematically as

$$Q = mc\Delta T \quad \text{Equation 1}$$

where m is the mass in kg, **c is the specific heat capacity in Joules/kg·C°**, and ΔT is the change in temperature in C°. When heat flows into a body and the temperature changes, the temperature will increase so ΔT and Q will be positive. Similarly, when heat flows out of a body and the temperature changes, ΔT and Q will be negative.

Specific heat capacity

When a quantity of heat Q, flows from one part of a system to another, the total internal energy of the system is unchanged. Thus Q represents the amount of energy simultaneously lost by one part of the system and gained by another. Given this it is easy to see why we use the term "flow" for this process.

In this investigation we want to predict the results of such an energy transfer and then confirm our prediction experimentally. If we place a mass of iron, m_i , at an initial temperature, T_{i0} , into our calorimeter and then add a mass of water, m_w , at an initial temperature, T_{w0} , we should find that heat flows from the warmer to the cooler body until thermal equilibrium is reached at the equilibrium temperature, T_f .

As observed above, Q_{net} for our (thermally isolated) system will be zero.

$$Q_I + Q_W = 0 \quad \text{Equation 2}$$

We can also write Equation 2 as

$$Q_I = -Q_W \quad \text{Equation 3}$$

For our system of iron and water we can combine Equations 1 and 3 giving us

$$m_I c_I \Delta T_I = -m_W c_W \Delta T_W \quad \text{Equation 4}$$

With Equation 4 we can predict a whole range of outcomes. We'll first try to predict the equilibrium temperature when known quantities of water and iron are mixed. Then we'll determine the specific heat capacity of an unknown material.

1. Determining the Equilibrium Temperature, T_f

We first need to expand Equation 4 to include the initial and final temperatures of both materials.

$$m_I c_I (T_{If} - T_{Io}) = -m_W c_W (T_{Wf} - T_{Wo}) \quad \text{Equation 5}$$

When the system is in thermal equilibrium, $T_{If} = T_{Wf}$, thus, replacing both of these with T_f ,

$$m_I c_I (T_f - T_{Io}) = -m_W c_W (T_f - T_{Wo}) \quad \text{Equation 6}$$

We can solve for T_f as follows.

$$\begin{aligned} m_I c_I T_f - m_I c_I T_{Io} &= -m_W c_W T_f + m_W c_W T_{Wo} \\ m_I c_I T_f + m_W c_W T_f &= m_I c_I T_{Io} + m_W c_W T_{Wo} \\ (m_I c_I + m_W c_W) T_f &= m_I c_I T_{Io} + m_W c_W T_{Wo} \\ T_f &= \frac{m_I c_I T_{Io} + m_W c_W T_{Wo}}{m_I c_I + m_W c_W} \quad \text{Equation 7} \end{aligned}$$

Using the masses and temperatures in Table 2, show the calculations to find the theoretical value for T_f here.

Now mix the two materials in your calorimeter to determine the final temperature experimentally. Record your experimental value for T_f in Table 2.

| Table 2 Mixing Hot Water with Cold Iron | | | |
|--|-----------|--|--|
| Specific Heat of Water, $c_W = 4186 \text{ J/kg}\cdot\text{C}^\circ$ | | | |
| Specific Heat of Iron, $c_I = 452 \text{ J/kg}\cdot\text{C}^\circ$ | | | |
| | Mass (kg) | Initial Temperature ($^\circ\text{C}$) | Final Temperature ($^\circ\text{C}$) |
| Water | .200 | 10 | |
| Iron | .800 | 80 | |

1. Calculate the percentage difference between your theoretical and experimental values for the equilibrium temperature.

% difference = _____

Show Calculations for % difference

2. Determining the specific heat capacity of an unknown material.

In addition to water, ice, and iron, the Dispens-O-matic can dispense a quantity of an unknown solid. Your task now is to determine the specific heat capacity of this mystery material. To do this you need to provide an appropriate **data table**, and **detailed calculations** beginning with the equation

$$m_M c_M \Delta T_M = -m_W c_W \Delta T_W$$

When you've determined c_M , use the tables in your textbook to **determine what material** you judge the mystery material to be. Indicate how accurate your result is by calculating the **percentage error**.

Use the space provided below to show your work.

Phase Changes and Latent Heat

So far we've only considered the case where heat flow results in a temperature change. Let's see what else can happen. In your calorimeter, mix together the materials listed in Table 3. The ice cubes are irregular so you may not get exactly 400 grams of ice. Watch the contents of the calorimeter as they approach equilibrium temperature.

| Table 3 Mixing Warm Water and Cold Ice | | | |
|---|-----------|--------------------------|------------------------|
| Specific Heat of Water, $c_w =$ | | 4186 J/kg·C° | |
| Specific Heat of Ice, $c_i =$ | | 2000 J/kg·C° | |
| | Mass (kg) | Initial Temperature (°C) | Final Temperature (°C) |
| Water | .400 | 90 | |
| Ice | .400 | -5 | |

Calculate $Q = mc\Delta T$ for each material.

$$Q_{\text{water}} = m_w c_w \Delta T_w = \underline{\hspace{2cm}} \text{ J}$$

$$Q_{\text{ice}} = m_i c_i \Delta T_i = \underline{\hspace{2cm}} \text{ J}$$

Clearly, $Q_1 \neq -Q_2$

We can see that the water lost a lot of energy. But the ice seems to have gained only about 7% of it. What gives? If you were watching carefully you saw that the ice slowly disappeared.

We stated earlier that the **absolute temperature** of a body is a measure of the **average internal energy per molecule**. And the **internal energy** of a body was defined as the sum of the molecular kinetic energy, the molecular potential energy, and other kinds of molecular energy. When the ice melts its molecules undergo a large increase in potential energy as they change from a rigid crystalline form to a liquid form. The warm water molecules are the source of this energy.

This change in potential energy is known as a **phase change** – the solid ice changed to liquid water. This significant amount of extra energy is accounted for by adding an extra term, $m_i L_f$, where L_f is the **latent heat of fusion**. The heat of fusion for ice is 33.5×10^4 J/kg.

A similar, but much larger amount of energy is required when water is converted from liquid to gaseous form. In that case, we have $m_w L_v$, where L_v is the **latent heat of vaporization**. We won't consider vaporization in this lab. The heat of vaporization for water is 22.6×10^5 J/kg.

Our ice warmed from -5°C to 0°C . It then underwent a phase change at 0°C , after which, in liquid form, it underwent a second temperature change of about 3.7°C . We can find the total energy involved as follows.

$Q = Q$ to heat the ice to 0°C + Q to melt the ice at 0°C + Q to heat melt water to 3.7°C .

$$Q = m_i c_i (T_{if} - T_{io}) + m_i L_{fi} + m_w c_w (T_{wf} - T_{wo})$$

$$Q = m_i c_i (0 - T_{io}) + m_i L_{fi} + m_w c_w (T_{wf} - 0)$$

$$Q = (.400 \text{ kg}) \left(2000 \frac{\text{J}}{\text{kg}\cdot\text{C}^\circ} \right) (0 - -5) \text{C}^\circ + (.400 \text{ kg}) \left(33.5 \times 10^4 \frac{\text{J}}{\text{kg}} \right) + (.400 \text{ kg}) \left(4186 \frac{\text{J}}{\text{kg}\cdot\text{C}^\circ} \right) (3.7 - 0) \text{C}^\circ$$

$$Q = 4000\text{J} + 1.34 \times 10^5\text{J} + 6195\text{J} = 1.44 \times 10^5\text{J}$$

This brings us more in line with our calculations for the heat lost by the 90°C water. And we can clearly see the large amount of energy involved in the phase change.

Before you continue with the final part of the lab, let's have a look at a graphical representation of this exchange of heat. On the data table below your lab apparatus there is a graph grid (right next to "you input this data.") Click on that grid to bring up a graph of temperature vs. heat. Follow the directions on the first frame of the animation.

I hope that gives you a clearer idea of what's been happening in the calorimeter. We'll finish by trying our hand at predicting that final temperature illustrated at the end of the animation.

Determining the final temperature of a mixture of ice and water.

Probably the most tedious calculations to make is the final temperature of a mixture. So let's give it a try and use your apparatus to validate your answer. Actually we'll do this in reverse just to give you a clearer picture of what to expect.

In your calorimeter combine at least 400 g of water hotter than 70°C with no more than 200 g of ice at a temperature of your choice. A **data table** is provided below to hold your data.

Starting from $Q_{lost\ by\ water} = -Q_{gained\ by\ the\ ice\ and\ melt\ water}$, write the full equation with all the individual terms. You should have four in all. Solve the equation for T_f showing all relevant steps. Substitute in all the data. Determine the theoretical value for T_f .

Finally compare your experimental and theoretical values for T_f by calculating a percentage error.

| Table 4 Mixing Warm Water and Cold Ice | | | | | |
|---|-----------|---------------|---------------------|----------------|----------------------|
| Specific Heat of Water, $c_w = 4186\text{ J/kg}\cdot\text{C}^\circ$ | | | | | |
| Specific Heat of Ice, $c_i = 2000\text{ J/kg}\cdot\text{C}^\circ$ | | | | | |
| Heat of Fusion of Ice, $L_{FI} = 3.35 \times 10^5\text{ J/kg}$ | | | | | |
| | Mass (kg) | T_{wo} (°C) | $T_f = T_{wf}$ (°C) | N/A | N/A |
| Water | | | | N/A | N/A |
| | Mass (kg) | T_{io} (°C) | T_{if} (°C) | T_{wio} (°C) | $T_f = T_{wif}$ (°C) |
| Ice | | | | | |

Note: T_{wi} refers to water from the melting of ice.

Show your equations and calculations in the space below

