VISUALIZING CHEMISTRY WITH INFRARED IMAGING

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Abstract: Almost all chemical processes release or absorb heat. The heat flow in a chemical system reflects the process it is undergoing. By showing the temperature distribution dynamically, infrared (IR) imaging provides a salient visualization of the process. This paper presents a set of simple experiments based on IR imaging to demonstrate its enormous potential for chemistry education. As the prices of IR cameras have plummeted recently, it is now time to consider bringing this powerful tool to every chemistry classroom.

INTRODUCTION

Scientists have long relied on powerful imaging techniques to see things invisible to the naked eye and thus advance science [1]. For example, microscopes and telescopes allow tiny and remote things to be observed, respectively.

Infrared (IR) imaging is a tool that shows the temperature distribution of a system based on detecting the invisible IR radiation it emits. Being a non-contact tool, it does not disturb the experimental system—an advantage over thermometers or temperature sensors that need to come into thermal contact with an object in order to measure its temperatures. The tool generates intuitive images in which different colors represent different temperature. As a picture is worth a thousand words, an IR camera is a perfect tool for teaching heat transfer [2-4].

More broadly, many physical, chemical, and biological processes that involve heat can be visualized by using IR imaging. In principle, *anything that leaves a trace of heat leaves a trace of itself under an IR camera*. To some extent, heat can be regarded as some kind of "IR ink" that renders a view of energy flow and reveals an invisible process. Deeply inside, as required by the Law of Conservation of Energy, any change of thermal energy (heat) must result in or result from the change of some kind of potential energy, which reflects some physical or chemical change in the system. Hence, by looking at the evolution of the thermal pattern of an experimental system under investigation, students can infer what is going on behind the images. This reasoning process often leads to a level as profound as molecular mechanisms, because heat and IR radiation ultimately originate from atoms, molecules, and their interactions. The purpose of this paper is to demonstrate, through a number of very simple chemistry visualization experiments, that IR imaging has the potential to become a versatile instrument in the classroom for students to see beyond perception, just like microscopes and telescopes for their corresponding subject matter.

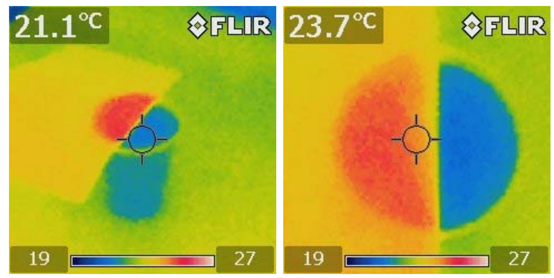
IR cameras used to be prohibitively expensive and difficult to use. Thanks to the breakthroughs in microsystem technologies in 1990s and the growing needs of home energy inspections and construction quality assurance using IR thermography in the past decade, IR cameras have become easy to use and their prices have plummeted [5, 6]. Basic versions of IR cameras that allow students to see heat flow in real time are now available for $1,500-$2,500 from FLIR (I5 and I7) and Fluke (TiS). The FLIR I5 was used in all the experiments presented in this paper. As easy to use as a typical digital camera, the I5 camera can automatically generate images of satisfactory quality with a temperature sensitivity of 0.1°C. Although the size of its microbolometer array is only 80 by 80 pixels (that is, nevertheless, 6,400 microsensors in just one camera!), it works very well for most lab bench experiments that do not need high resolution.

This paper presents a set of hands-on and minds-on experiments devised to tap the educational potential of IR imaging in fostering inquiry-based learning in chemistry education. All these experiments are very easy to do. But they all come with one or more science puzzles that can stimulate students' interest and engage them to explore more deeply in science. As the price of IR cameras continues to drop, it is now time to introduce this fascinating tool to chemistry teachers.

SEEING EVAPORATION, CONDENSATION, AND LATENT HEAT

This experiment involves just two identical plastic cups and a piece of paper (and, of course, an IR camera). Put the two cups side by side on a table. Fill one of the cups with room temperature tap water and place the paper on top of them in such a way that it covers approximately half of the water surface (the other cup is really just used to support the piece of paper). Make sure that the cup is nearly full and the paper is not wetted or soaked by water.

Use an IR camera to take a look immediately after the paper is placed. Figure 1 shows what we observed. [Since these are the first IR images, let us briefly explain how to read them. In an IR image, the number at the upper-left corner is the temperature of the spot to which the crosshair points (it acts like an IR thermometer). The numbers at the bottom are the lower and upper bounds of the temperature detected by the IR camera. The colors represent the temperature: red means hot, blue means cold, and any color inbetween means temperature inbetween, as shown by the rainbow color bar at the bottom. All the IR images in this paper were taken with the emissivity parameter of the IR camera set to 0.80.]



(a) (b)

Figure 1. An IR image of a piece of paper half covering the top of a cup of water shows the cooling effect due to evaporation at the water surface and the warming effect due to condensation on the paper. (a) Far side view. (b) Near top view. The images were taken shortly after the paper was placed.

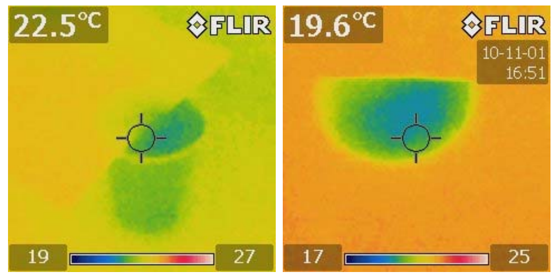
Figure 1 shows that the part of the paper above water was warmer than the rest of it, which was approximately at room temperature. In contrast, the entire cup appeared to be cooler.

Most students may be able to reason that the cup should be cooler because water constantly evaporated from it to take away thermal energy. But exactly how did evaporation at the surface cool down the *entire* cup all the way to the bottom? This is actually a complicated process. First, water evaporated from the surface and lowered the temperature of the top layer. The cooler layer then sank and, on its way to the bottom, cooled the water below. Once the water became cooler than room temperature, energy from the environment was transferred in through the sidewall of the cup or the air above it to warm it up a bit and prevent the temperature from decreasing further. This process ran continuously to keep the whole cup slightly cooler than the environment as water constantly evaporated from the surface. One can confirm this effect of evaporative cooling by sealing the cup with plastic wrap to stop the evaporation. A sealed cup of water will vanish in an IR image after some time, meaning that the temperature of the whole cup becomes identical to the ambient temperature. (This can be an activity students can do, too.)

The natural thermal convection in an open cup of water is hardly observable, however, because the amount of heat lost through evaporation is small and the process occurs slowly (for a 250ml cup full of water, the evaporation rate is about 6 grams per day in a typical office environment in New England). Later in this paper, we will show a convective flow with a more dramatic setup.

Now, the interesting question is: why was the part of the paper above the water warmer? This result, which would probably surprise many students, shows condensation at work. It was the latent heat of the condensation of water vapor onto the underside of the paper that warmed up that side and then the energy was conducted through the paper. The amount of conducted heat was large enough to show up in the IR image—the area registered about 1°C warmer than the ambient temperature.

This simple experiment is a visual demonstration of the latent heat of condensation. But there are more puzzles that teachers can challenge students to solve. If we leave the paper on the cup for a long time, say a couple of hours, and then come back to look at its IR image, Figure 2a shows what we will see. The warm area vanished from the IR image, meaning that the entire paper was almost at room temperature. Why did the condensation heating stop? The answer lies in that the paper surface can only take a limited amount of condensate. At the same time water vapor condensed to the surface and coated it with a thin layer of liquid water, water molecules in the layer had to evaporate as well. The layer could not grow thicker after the rate of evaporation balanced that of condensation. When the equilibrium was reached, there was no more net heat released to warm up the paper, which then stayed at the ambient temperature. In the meantime, the evaporation cooling at the uncovered half of the water surface never ceased because that part was open to the entire environment and water vapor could diffuse away. The acute reader may notice that the entire cup appears to be warmer in Figure 2a than in Figure 1a. This is because the evaporation surface of the cup was cut half and so was the overall evaporative cooling effect. In the extreme case when the entire cup is covered, the evaporative cooling effect will stop and the whole cup will vanish in the IR view (meaning that it will be at room temperature).



(a) (b)

Figure 2. (a) An IR image of a piece of paper half covering the top of a cup of water taken after the paper has been placed for a few hours. (b) An IR image of the paper taken immediately after it was removed from the top of the cup.

How can we be sure that there was water condensate on the paper? We probably cannot feel the moisture by touching the paper because the amount of condensed water is so tiny. But if we remove the paper from the cup and then immediately view the paper through an IR camera, we will see the evaporative cooling effect, as is shown in Figure 2b. The bluish semicircular mark with a greenish edge in the image was the trace left by the condensate, which quickly faded away as water evaporated from the paper.

We hope you are now reasonably inspired by the power of IR imaging demonstrated through this first experiment. Although very easy to do, there is a considerable depth of science embodied in the experiment, as analyzed above. With an IR camera, those science concepts can be visualized and discovered in an unprecedentedly straightforward way. In the following, we will show you more examples.

SEEING THE HEAT OF SOLUTION

Prepare some room temperature tap water and three clean identical plastic cups for making three liquids: freshwater, unsaturated saltwater, and saturated saltwater. Depending on the size of your cups, add some salt (e.g. 20g for a 250 ml cup) to one of them and a lot of salt (e.g. 60g for a 250 ml cup) to another. Fill all three cups up with room temperature tap water and then immediately aim an IR camera at them. Figure 3a shows that all three cups appeared to be cooler than the environment. The temperature of the cup of freshwater was the same everywhere, whereas the lower parts of the other two cups were cooler than the upper parts, indicating that salt at the bottoms were dissolving and the process was endothermic. Because the middle cup had less salt than the right one, the cooling effect was weaker.

Besides salt, students can try different substances such as baking soda and sugar. For example, one can compare the heats of solution of table salt and baking soda by adding the same weight of each to the same amount of water in two cups. The heats of solution for table salt ($NaCl$) and baking soda ($NaHCO_3$) are

3.9 and 15.6 kJ/mol, respectively. IR imaging clearly shows that the heat of solution of baking soda is significantly higher than that of table salt (Figure 3b).

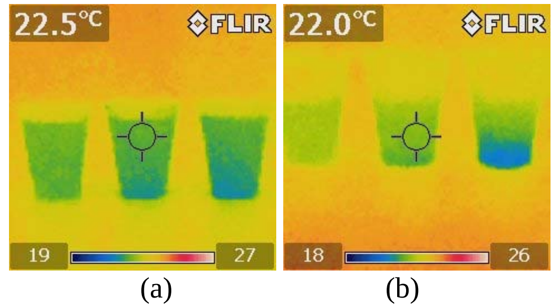


Figure 3. (a) An IR image of three cups taken just after water was added. There was no salt in the left cup, 20g of salt in the middle one, and 60g in the right one. The IR image shows the salt at the bottoms of the middle and right cup was dissolving. (b) An IR image of three cups taken just after water was added. There was no solute in the left cup, 20g of salt in the middle one, and 20g of baking soda in the right one. In both (a) and (b), there is no emissivity issue as the images show the temperature of the thin plastic sidewalls of three identical cups that have the same emissivity.

DISCOVERING A TEMPERATURE GRADIENT

The previous experiment on the heat of solution seems interesting but not surprising. Now add more salt and baking soda to the cups to ensure that the solutions are saturated and leave the cups on the lab bench for a long time, say 24 hours, to allow them to settle. Do not place them under strong light, near a furnace, or in airflow. When you come back to look at them through an IR camera, you should see something baffling.

Figure 4a shows what we observed. The results for the two cups of solutions show that the temperature at the bottoms was about 0.4-0.5°C *higher* than the temperature at the tops. And we found that this temperature gradient lasted for a long time—at least a couple of weeks. We have also confirmed the results by using a non-IR method with a fast-response surface temperature sensor from Vernier [7] in cups made of different materials (e.g. paper or glass). We concluded that this effect is not an artifact of IR imaging and it has nothing to do with what the cups are made of.

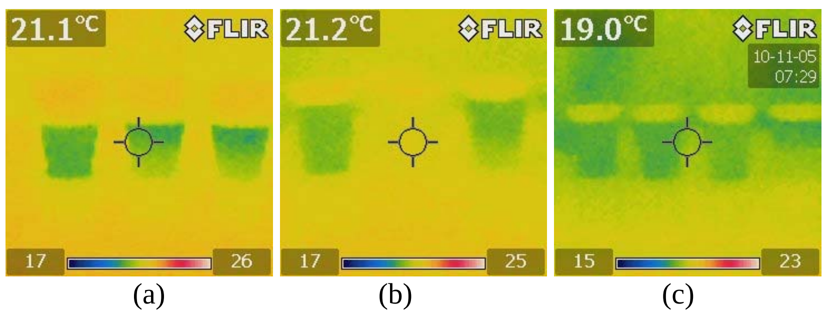


Figure 4. (a) An IR image of the three cups of liquid—freshwater (left), saturated saltwater (center), and saturated baking soda water (right)—shows that the lower parts of the solutions were constantly warmer than the upper parts. (More greenish means cooler.) (b) An IR image of the three cups after the middle one was sealed shows that it reached a thermal equilibrium with the environment. (c) An IR image of four cups of liquids shows that a significant temperature gradient was only present in the saturated solution. From left to right: freshwater, low-concentration saltwater, medium-concentration saltwater, and saturated saltwater.

Why should there be a temperature gradient in the saturated solutions? We know that there is natural convection in pure water—warmer water tends to rise and cooler water tends to sink—that should eventually eliminate any temperature difference, just like what the IR image of the freshwater cup on the left in Figure 4a shows. But something stopped the thermal convection in the two cups that contain a saturated solution. Exactly what is the chemical force that drives this persistent temperature gradient? We searched the literature and did not feel that the physical origin of this phenomenon was well understood (or even reported before). So we decided to do more investigations to figure it out ourselves.

In the first investigation, we sealed the cup of saltwater (the middle one) with a plastic wrap and gave it 24 hours to settle. Figure 4b shows an IR image of the three cups after that. The complete disappearance of the sealed cup in the IR image means that the temperature in the saltwater became the same everywhere and it was equal to the ambient temperature. That is to say, when evaporation stopped, not only did the evaporative cooling cease but also was the temperature gradient gone (or reduced to less than 0.1°C—the smallest temperature difference our IR camera is capable of detecting). The entire cup reached a thermal equilibrium with the environment. This underpins the importance of evaporation but does not explain why the bottom of an open cup of saturated saltwater was always warmer than the top.

In the second investigation, we compared four cups of liquids: freshwater, low-concentration saltwater, medium-concentration saltwater, and saturated saltwater. After sitting on the lab bench for a couple of days, their IR image shown in Figure 4c suggests that a significant temperature gradient was only present in the saturated saltwater. The cups of unsaturated saltwater had approximately the same temperature as freshwater, regardless of the salt concentration. The results were confirmed independently by using a Vernier temperature sensor to measure the temperature at difference depths.

How do we explain all these results? There are two candidate theories.

<u>The Crystallization Theory</u>. Since water molecules are constantly evaporating from the surface of the saltwater, a corresponding amount of ions must return to the solid form at the same time—because a reduced amount of water in a saturated solution in the cup cannot accomodate them any more. When the ions precipitate at the bottom, they adhere to the surface of a crystal and the process releases heat to surrounding water molecules. This heat of crystallization at the bottom, together with the evaporative cooling effect at the surface, creates the temperature gradient. The entire process runs continually across the solution because of the diffusion of water molecules and ions driven by their corresponding concentration gradients. This theory can explain the experimental results. In the first investigation, the temperature gradient disappeared when the cup was sealed, because that stopped the evaporation at the surface and, therefore, the net crystallization at the bottom. Without the decrease of water molecules in the solution, the crystallization and dissolving processes at the bottom reached equilibrium. No net growth of salt crystals would occur and no net heat would be released. In the second investigation, no significant temperature gradient could be observed in an unsaturated solution because there was no crystallization process at the bottom. Lastly, this temperature gradient persisted for a long time because this process would continue until all the water molecules in the cup evaporate.

<u>The Buoyancy Theory</u>. There is a salinity gradient in the cup—the saltwater contains more salt at the bottom than at the top (which can be confirmed with a salinity sensor). While the top layer of the water is cooled by evaporation, it cannot sink because the saltier water below it provides greater buoyancy that counterbalances the convective force. The formation of the salinity gradient in a cup of saltwater can be explained with the Boltzmann distribution, similar to the explanation of the barometric formula that describes how density changes with respect to altitude. This theory can explain the result of the first investigation as well. When the cup was sealed, evaporation stopped. Although the cooler water in the upper part could not sink through convection, its temperature would eventually become the same as everywhere in the cup due to molecular diffusion and heat conduction. It has trouble explaining the result of the second investigation, however. But this does not necessary invalidate this theory, as our further studies show that the salinity gradient in a cup of unsaturated saltwater is actually much smaller than that in a cup of saturated saltwater. As these studies fall out of the scope of this paper, we will not discuss them further.

It seems to the author both theories are right to some extent. The truth may be that the two mechanisms together provide the thermodynamic driving forces to maintain the mysterious temperature gradient. What is important in this experiment is that this subtle, puzzling effect might never have been noticed in the classroom without an IR camera. Having students discover it themselves and challenging them to solve the puzzle through further investigations can provide plenty of opportunities of inquiry and engagement to learn the science more deeply. In the following, we will show one of the experiments we devised to help solve this puzzle.

SEEING THE EFFECT OF SALT ON NATURAL CONVECTION

The thermal convection in the previous experiments was not evident because the temperature difference was very small. To better understand thermal convection in freshwater and saltwater, let us do the following simple experiment in which the temperature difference is greatly increased.

Prepare a cup of freshwater and a cup of saturated saltwater and add an ice cube to each of them. Figure 5 shows a series of IR images taken after that. The results clearly indicate that the thermal convection that happened in the freshwater to cool down the entire cup somehow did not happen in the saltwater. This experiment provides evidence that supports the Buoyancy Theory.

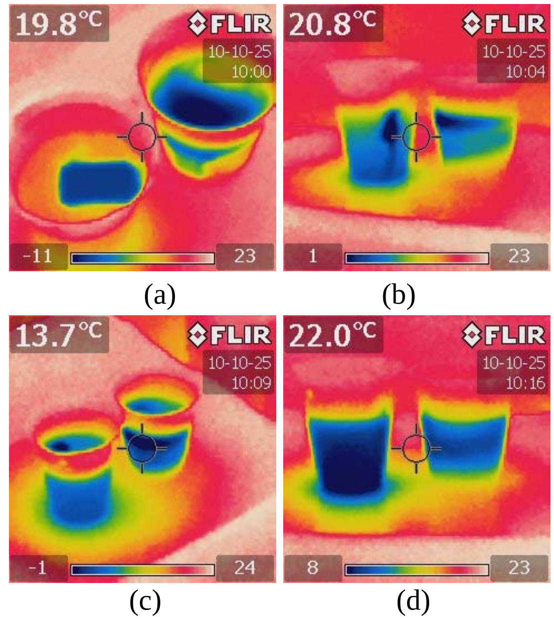
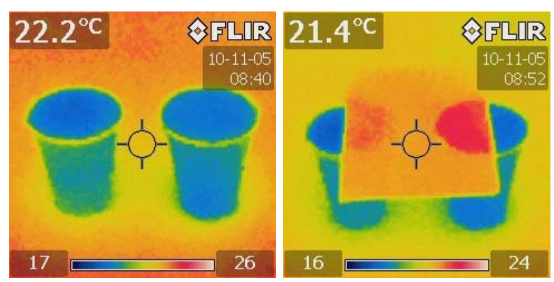


Figure 5. (a) An IR image right after an ice cube was added to a cup of freshwater (left) and a cup of saturated saltwater (right). (b) An IR image taken after four minutes showing significant convection in the freshwater cup. (c) An IR image taken after nine minutes showing the tabletop was cooled significantly near the freshwater cup. (d) An IR image taken after sixteen minutes showing that the bottom of the freshwater cup became cooler than the top whereas the bottom of the saltwater cup remained warmer than the top.

SEEING VAPOR PRESSURE LOWERING

The vapor pressure lowering is an effect in which the water vapor pressure above saltwater is lower than that above freshwater. This is more generally described by Raoult's Law, which states that the vapor pressure of an ideal solution depends on the vapor pressure of each chemical component and the mole fraction of the component present in the solution.

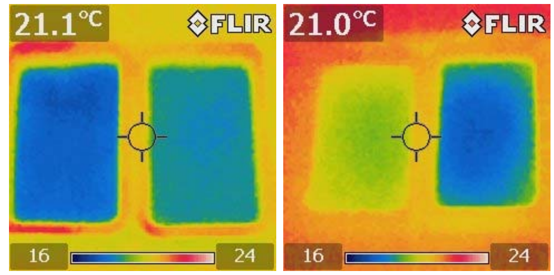
Let us use IR imaging to visualize this effect. Prepare a cup of saltwater and a cup of freshwater and place them side by side as shown in Figure 6a. Put a piece of paper on top of them and compare the condensation heating effect discussed earlier. Figure 6b shows that the warming effect on the paper was more significant above the freshwater cup than above the saltwater cup, suggesting that the humidity and, therefore, the vapor pressure above saltwater is lower than that above freshwater. Note that the red marks in the IR image of the paper would not last long—the paper would eventually reach thermal equilibrium and they would disappear, just as Figure 2a shows.



(a) (b)

Figure 6. (a) An IR image of a cup of saltwater (left) and a cup of freshwater (right). Note that the left cup appeared to be slightly warmer than the right one, suggesting that evaporation was depressed in the saltwater cup. (b) An IR image taken shortly after a piece of paper was placed on top of the two cups shows less warming on the paper above the saltwater side, providing strong visual evidence that less evaporation from the saltwater surface.

How do we capture a stable IR image of vapor pressure lowering that will last longer? Note that in Figure 6 the surfaces of saltwater and freshwater appeared to be nearly as cool. One might expect to see a detectable difference in temperature resulting from the difference of evaporation rates. This can be demonstrated using shallower containers. Prepare two shallow plastic containers (10cm×5cm×1cm in our experiments). Add plenty of salt to one of them and then fill both with room temperature tap water. Figure 7a shows an IR image just after water was added. The image shows that the system absorbed heat while salt was being dissolved, something we have seen previously at the bottom of a cup of water after we added salt to it. This effect now shows up on the surface of the saltwater because the water was very shallow and energy could be conducted through it in a short time.



(a) (b)

Figure 7. (a) An IR image taken right after water was added to the shallow containers with salt (left) and without salt (right). (b) An IR image taken two hours later.

Let the containers sit for a few hours and then take another IR image. Figure 7b shows the result. The saltwater surface appeared to be significantly warmer than the freshwater surface. Besides the effect of slower evaporation at the surface of saltwater, another possible contribution to the warming effect comes from the heat of salt crystallization, as explained in the Crystallization Theory in a previous section, in the case when the solution is saturated.

CONCLUSIONS

By allowing students to see phenomena that would otherwise be invisible, IR imaging provides a vivid real-time visualization tool for chemistry education. All the experiments demonstrated in this paper are very easy to implement. Yet they reveal very profound science that might cause even experts to scratch their heads. What a cup of saltwater can teach under an IR camera is surprisingly plentiful.

The reader may be wondering if IR imaging will be useful in teaching other topics beyond the above examples. We can imagine that IR imaging could be used to visualize other chemical processes such as reactions. In many cases, the trick for designing successful IR activities is to come up with clever setups in which the surface heating or cooling that an IR camera will pick up can reflect the chemical processes in the bulk (such as placing a piece of paper atop a cup of water). Our future work in visualizing chemistry with IR imaging will showcase more experiments designed in this way.

From a pedagogical point of view, IR imaging has great potential to enable novel inquiry designs to transform students' learning experience with chemistry. It significantly lowers the technical barrier of experimental skills that would have been needed to reach the same depth of science and the same level of inquiry. For example, a novice learner, if just given a temperature sensor, might never be able to discover the mysterious temperature gradient in a cup of saltwater or the condensation heating effect on a piece of paper above water, because the work would be more laborious and require higher analytical skill to connect the dots. Compared with a temperature sensor that presents data in the form of numbers or graphs, IR imaging presents colorful, intuitive visualizations that may be more compelling and comprehensible to students. With all this exceptional educational potential, the power of IR imaging for teaching and learning chemistry should not continue to go unutilized.

The current cost of an economic version of IR camera may still be considered as a barrier. But if enough number of science teachers start to appreciate the value of this tool and every school starts to buy a few for their labs, a sizable education market will emerge and the prices of the IR cameras will be driven down further, which will in turn benefit both the industry users and the educational users. By the time this paper was being written, there was still no educational discount from the manufacturers and vendors, citing insufficient orders. This paper, along with our current and future work that will spearhead creative applications of IR imaging to the education of chemistry, physics [2], biology, earth science, and engineering, will hopefully pave the road for IR imaging to reach science classrooms.

Acknowledgment

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