Don’t Flip Your Lid
Comparing Intermolecular Forces

About this Lesson
This lesson is a short experiment which allows students the opportunity to apply knowledge of bonding and intermolecular forces to a practical situation regarding melting points.

This lesson is included in the LTF Chemistry Module 4.

Objectives
Students will:

- Use their knowledge of intermolecular forces (IMF) to explain relative differences in melting points.

Level
Chemistry

Common Core State Standards for Science Content
LTF Science lessons will be aligned with the next generation of multi-state science standards that are currently in development. These standards are said to be developed around the anchor document, A Framework for K–12 Science Education, which was produced by the National Research Council. Where applicable, the LTF Science lessons are also aligned to the Common Core Standards for Mathematical Content as well as the Common Core Literacy Standards for Science and Technical Subjects.

<table>
<thead>
<tr>
<th>Code</th>
<th>Standard</th>
<th>Level of Thinking</th>
<th>Depth of Knowledge</th>
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</thead>
<tbody>
<tr>
<td>(LITERACY) RST.9-10.3</td>
<td>Follow precisely a multistep procedure when carrying out experiments, taking measurements, or performing technical tasks, attending to special cases or exceptions defined in the text.</td>
<td>Apply</td>
<td>II</td>
</tr>
<tr>
<td>(LITERACY) RST.9-10.7</td>
<td>Translate quantitative or technical information expressed in words in a text into visual form (e.g., a table or chart) and translate information expressed visually or mathematically (e.g., in an equation) into words.</td>
<td>Apply</td>
<td>II</td>
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<tr>
<td>(LITERACY) W.1</td>
<td>Write arguments to support claims in an analysis of substantive topics or texts, using valid reasoning and relevant and sufficient evidence.</td>
<td>Apply</td>
<td>II</td>
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<tr>
<td>(LITERACY) W.4</td>
<td>Produce clear and coherent writing in which the development, organization, and style are appropriate to task, purpose, and audience.</td>
<td>Apply</td>
<td>II</td>
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<tr>
<td>(LITERACY) RST.9-10.9</td>
<td>Assess the extent to which the reasoning and evidence in a text support the author’s claim or a recommendation for solving a scientific or technical problem.</td>
<td>Apply</td>
<td>II</td>
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Connections to AP*

AP Chemistry:
I. Structure of Matter
   B. Chemical bonding
      1. Binding forces
         a. Types: ionic, covalent, metallic, hydrogen bonding, van der Waals (including London dispersion forces)
         b. Relationships to states, structure, and properties of matter
         c. Polarity of bonds, electronegativities

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Materials and Resources

Each lab group will need the following:
- aprons
- Bunsen burner
- can lid
- dextrose
- goggles
- iodine, solid
- marker, Sharpie®
- paper towels
- paraffin wax
- ring stand and ring clamp
- salt
- wood splint

Assessments

The following types of formative assessments are embedded in this lesson:
- Assessment of prior knowledge regarding bonding and IMF.
- Visual assessment of laboratory design.

The following additional assessments are located on the LTF website:
- Chemistry Assessment: Bonding
- AP Style Free Response
Teaching Suggestions
This is a simple activity that takes little time. However, a good explanation requires students to have a relatively thorough understanding of intermolecular forces and the role they play in physical properties such as melting point.

The four compounds tested are NaCl (an ionic compound), paraffin (a medium sized nonpolar compound), glucose (a medium sized polar compound), and iodine (a larger nonpolar molecule). They are placed on the outer rim of the can lid and the Bunsen burner is moved under the center of the lid. The compounds melt in this order: paraffin, iodine, glucose. The sodium chloride does not melt.

*Note:* The use of the terms IMF, van der Waals, London forces and dispersion forces varies widely among textbooks and resources. Past convention was to use the term van der Waals as a generic term referring to any type of intermolecular attraction. The recent trend in the literature, however, is to use van der Waals to indicate the forces developed due to the polarization of the electron cloud. In other words, it is synonymous with London dispersion forces. Conventions change as the years go by so be flexible with the term van der Waals in your classroom and alert students to the variety of meanings they may encounter.
Answer Key

Data and Observations
The compounds melt in this order: paraffin, iodine, glucose. The sodium chloride does not melt.

Analysis
Student answers will vary. However, a good answer will include the following points:

- Ionic bonds are very strong. They are electrostatic and work in all directions. Therefore, the sodium chloride does not melt under these lab conditions. Its melting point is 801°C.

- The other issues involved concern whether the molecule is polar or nonpolar and the number of electrons the molecule has. All the molecules have electrons, and therefore have London forces. However, glucose is also polar and can form hydrogen bonds. Therefore, it has stronger IMFs and a greater melting point.

  Between the remaining two molecules, paraffin and iodine, both have only London forces. However, iodine has a more polarizable electron cloud. Therefore, it has stronger intermolecular forces and a greater melting point.

- Be sure that students do not think that the more massive molecule has stronger IMFs. The problem with this argument is that it confuses gravitational attraction with London forces. Gravitational attraction is a function of the mass of the particle but it is very weak, and does not really play a part in the attraction between molecules.

  London forces are the electrostatic forces that arise from the distortion of the electron cloud. The more electrons the molecule has—and the more loosely the electrons are held—the stronger the IMFs. You might think of this as the “squishiness” of the electron cloud. This is referred to as polarizability.
Conclusion Questions

1. Sucrose has a greater melting point. It has more electrons, and therefore its electron cloud is more polarizable and results in stronger IMFs.

2. H₂S and H₂O are both bent molecules and are polar. However, H₂O has its hydrogen bonded directly to a small, electronegative atom—oxygen. Therefore, intermolecular hydrogen bonding can occur between adjacent water molecules.

   ![Hydrogen bond](image)

   The sulfur in H₂S is not a highly electronegative atom, and thus H₂S does not allow for intermolecular hydrogen bonding. This molecule is polar and exhibits dipole-dipole interactions.

3. All of the halogens are nonpolar diatomic molecules. All of the halogens have only London forces.

   Iodine is a solid because it has the greatest melting point. It has more electrons than the other halogens, and therefore its electron cloud is more polarizable and results in stronger IMFs than bromine, chlorine, or fluorine.

   Bromine is a liquid, which means its melting point is greater than fluorine or chlorine. It has fewer electrons than iodine but more electrons than fluorine or chlorine. Bromine's electron cloud is more polarizable than that of fluorine or chlorine, and therefore it has stronger IMFs than fluorine or chlorine.

   Fluorine and chlorine are small molecules with very few electrons. These molecules must be cooled considerably and thus slowed down before their IMFs are strong enough to draw them close together to liquefy, or subsequently solidify.
The forces that hold one molecule to another molecule are referred to as *intermolecular forces (IMFs)*. These forces arise from unequal distribution of the electrons in the molecule and the electrostatic attraction between oppositely charged portions of molecules.

*Van der Waals forces* are a function of the number of electrons in a given molecule and how tightly those electrons are held. Let us assume that the molecule involved is nonpolar. A good example would be O₂. Pretend that the molecule is all alone in the universe. If that were the case, the electrons in the molecule would be perfectly symmetrical. However, the molecule is not really alone. It is surrounded by other molecules that are constantly colliding with it. When these collisions occur, the electron cloud around the molecule is distorted. This produces a momentary induced dipole within the molecule. The amount of distortion of the electron cloud is referred to as polarizability. Since the molecule now has a positive side and a negative side, it can be attracted to other molecules. This attractive force is called a London force or a dispersion force. Since all molecules have electrons, all molecules have London forces. These forces range from 5–40 kJ/mol.

Some molecules are naturally polar. Therefore, in addition to dispersion forces, they can also have a permanent dipole which attracts other polar molecules (either induced or permanent).

A third type of intermolecular force is *hydrogen bonding*. When hydrogen is covalently bonded to a small electronegative atom like nitrogen, oxygen, or fluorine, the electron cloud on the hydrogen is very distorted and pulled toward the electronegative atom. Since hydrogen has no inner core electrons, the positive nuclear charge is somewhat exposed. This sets up the potential for a reasonably strong attraction between this hydrogen and an electronegative atom in another molecule. Hydrogen bonds are significant in determining such factors as the high boiling point of water, solubility of acetone in water, and the shape and structure of proteins and DNA.

Ionic bonds are formed between charged particles. Ionic compounds do not form molecules, but rather are held together in a crystalline structure by electrostatic attraction. These electrostatic attractions are strong and are omnidirectional.
PURPOSE
In this laboratory activity you will determine the relative melting points of four substances, paraffin, NaCl (table salt), C₆H₁₂O₆ (glucose), and iodine, I₂. Paraffin is a medium size nonpolar molecule, NaCl is an ionic compound. Glucose is a medium size polar molecule. Iodine is a large nonpolar molecule.

MATERIALS

*Each lab group will need the following:*
- aprons
- Bunsen burner
- can lid
- dextrose
- goggles
- iodine, solid
- marker, Sharpie®
- paper towels
- paraffin wax
- ring stand and ring clamp
- salt
- wood splint

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**Safety Alert**
1. Wear your goggles.
2. Do not inhale the iodine vapors.
3. Allow the can lid to cool before disposal.
PROCEDURE
1. In the space marked hypothesis on your student answer page, write a statement predicting in which order the salt, paraffin, iodine, and glucose will melt.

2. Turn the can lid over so that the grooves around the edge form a trough. Use the grease pencil or marker to divide the lid into four quadrants as shown in Figure 1.

3. Use wooden splints to transfer small samples of each of the four compounds to the outermost groove of the can lid. You only need a very small sample—just enough to be able to see the compound. For the iodine, use only one crystal.

4. Carefully place the can lid onto the ring of the ring stand as shown in Figure 2. Light the Bunsen burner and place it so that the flame is directly under the middle of the can lid.

5. Watch carefully as the compounds melt. You will need to record the relative order of melting. As soon as three of the compounds melt, turn off the burner.

6. Allow the can lid to cool and dispose of it in a trash can.

7. Answer the conclusion questions on your student answer page.
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HYPOTHESIS

DATA AND OBSERVATIONS

Record the relative order of melting for the four compounds.

ANALYSIS

From what you know about intermolecular forces, explain the relative order of melting points.
CONCLUSION QUESTIONS

1. Glucose, C$_6$H$_{12}$O$_6$, and sucrose, C$_{12}$H$_{22}$O$_{11}$, are both sugars with similar polarities. Which do you think would have the higher melting point? Justify your answer.

2. Hydrogen sulfide is a gas at room temperature, while water is a liquid, yet hydrogen sulfide has more electrons than water. Explain this anomaly.

3. Consider the halogens at room temperature and 1 atmosphere of pressure. Why are fluorine and chlorine gases at room temperature, while bromine is a liquid and iodine a solid?