



# Don't Flip Your Lid

## Comparing Intermolecular Forces

### MATERIALS AND RESOURCES

#### EACH GROUP

aprons	marker, Sharpie®
goggles	dextrose
hot plate	paraffin wax
paper towels	sucrose
can lid	wood splint
	salt

### ABOUT THIS LESSON

This lesson is a short experiment that asks students to apply concepts of bonding and intermolecular forces to justify the ranking of experimentally determined melting points for four solids.

### OBJECTIVES

Students will:

- Use relative melting point data to apply and solidify their knowledge of intermolecular forces (IMF)

### LEVEL

Chemistry

**COMMON CORE STATE STANDARDS**

**(LITERACY) RST.9-10.3**

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.

**(LITERACY) RST.9-10.7**

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.

**(LITERACY) RST.9-10.9**

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.

**(LITERACY) W.1**

Write arguments to support claims in an analysis of substantive topics or texts, using valid reasoning and relevant and sufficient evidence.

**(LITERACY) W.4**

Produce clear and coherent writing in which the development, organization, and style are appropriate to task, purpose, and audience.

**NEXT GENERATION SCIENCE STANDARDS**



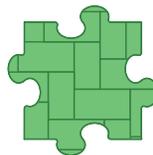
ASKING QUESTIONS  
DEFINING PROBLEMS



ANALYZING AND  
INTERPRETING DATA



CONSTRUCTING EXPLANATIONS  
DESIGNING SOLUTIONS



PATTERNS



STRUCTURE  
AND FUNCTION



PS1: MATTER



PS2: FORCES  
AND INTERACTION



PS3: ENERGY

**ACKNOWLEDGEMENTS**

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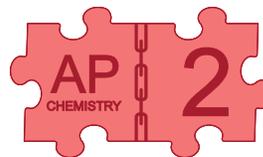
**ASSESSMENTS**

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 our website:

- Chemistry Assessment: Bonding
- AP Style Free Response

**CONNECTIONS TO AP\***

- B.1** London dispersion forces are attractive forces present between all atoms and molecules. London dispersion forces are often the strongest net intermolecular force between large molecules.
- B.2** Dipole forces result from the attraction among the positive ends and negative ends of polar molecules. Hydrogen bonding is a strong type of dipole-dipole force.
- B.3** Intermolecular forces play a key role in determining the properties of substances, including biological structures and interactions.
- C.1** In covalent bonding, electrons are shared between the nuclei of two atoms to form a molecule or polyatomic ion. Electronegativity differences between the two atoms account for the distribution of the shared electrons and the polarity of the bond.
- C.2** Ionic bonding results from the net attraction between oppositely charged ions, closely packed together in a crystal lattice.

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## TEACHING SUGGESTIONS

In this lesson, students will compare the relative melting points of four solids to create the context for a teacher-facilitated discussion on intermolecular forces. The compounds tested are sodium chloride (an ionic compound), paraffin (a large nonpolar molecule), dextrose (a medium-sized polar molecule), and sucrose (a larger polar molecule).

Small samples of each substance are placed on the outer rim of a can lid and heated slowly and uniformly on the surface of a hot plate until the first three substances melt. At that point, students will know the relative ranking of melting point for all four substances. The compounds should melt in this order: paraffin, dextrose, and sucrose. The sodium chloride does not melt.

Ideally, this investigation is done after students have explored bonding and polarity. This lesson models the use of a simple laboratory activity as a means of establishing a common experience from which the concept of intermolecular forces can be introduced. The intent of this lesson is best executed when students have not yet had instruction on intermolecular forces, as this activity is not designed to be a confirmation lab.

Students may have already encountered Coulomb's law in discussions on atomic theory and bonding. The background section serves as an independent review for students and brings the concept of coulombic interactions to the forefront of their thinking as they approach the lab. Begin this lesson by asking students to "actively read" or "read for understanding" the introductory text of the lab. Remind students of strategies for interacting with text that have been practiced in class. Such strategies may include highlighting important phrases, circling domain-specific vocabulary, annotating paragraph summaries, and identifying the main idea of the text.

It is your discretion about whether the Pre-Lab Exercises are provided to students before or after an initial read-through of the text. This decision should be based on how adept students are at independently applying reading skills. If students are answering pre-lab questions without purposeful interaction with the text, as evidenced by their mark-ups, then withholding the lab questions until after at least one read-through is recommended. The background section can also be assigned to students prior to class, in which case part of the assessment for this task should be looking for evidence that students interacted with the text as part of their question-answering process.

Before the lab, allow time for sharing of student responses to the Pre-Lab Exercises after students have had a chance to answer them independently or collaboratively with a partner. Note that Question 2 and Question 3 ask for responses that are best supported by a *Claim-Evidence-Reasoning (CER) framework*. It is advisable to prepare example student responses to these questions as a model of what constitutes a proficient answer to these questions.

The following questions may be useful while fielding the whole-group sharing of student responses for Question 2 and Question 3:

- Was the claim clear in that response? Can someone restate the claim the student made?
- What was the objective evidence used in the response?
- After asking the student to reread the response, ask: Does the response include clear reasoning? If yes, does the reasoning connect back to the claim? If no, what reasoning could be added to complete this response?

**TEACHING SUGGESTIONS (CONTINUED)**

In the discussion around student responses, the following content points are most important for students to possess clarity:

- Attractions happen between opposite charges.
- Repulsions happen between like charges.
- The force of the attraction is inversely related to the square of the distance between the charges. Two opposite charges experience a much greater force of attraction at smaller distances. Two like charges experience a much greater force of repulsion at smaller distances. For struggling learners, demonstrating this concept with two like poles of a magnet should help make this abstract idea visual and more understandable.
- At the melting point, particles that are in a specific solid structural arrangement gain more motional freedom without changing chemical composition.

Following the discussion of the concepts addressed in the background section and Pre-Lab Exercises inform the students that they will be investigating the relative melting points of four white solids. Small amounts of each substance should be set out on the lab tables with wood splints. If using hot plates, the hot plate should be unplugged at the start of the investigation.

Ask students to read through the entire set of procedures before beginning, and monitor students to ensure they fill out the prediction section. It is acceptable if students cannot clearly articulate the reasons behind their prediction. The intent of that step is to get them thinking of factors that could contribute to the melting point, as this creates an authentic “need-to-know,” setting the stage for the information the teacher will provide later in the Analysis section.

As students perform the investigation, prompt them to watch very carefully and make sure the students do not mix up their white solids. If technology is permitted in the classroom, some students may want to record this process so they can verify the order of melting if it happens too quickly. The relative order of melting point should be recorded under Data and Observations while students are at their stations.

After students have completed the lab and cleaned their stations, they may complete Question 1 through Question 3 of the Analysis section. Question 1 relates what happened in the lab to a new term, *intermolecular forces*, considering the concept of melting point addressed in Question 5 of the Pre-Lab. Question 2 requires students to apply prior knowledge of bonding and polarity to fill in the table.

## TEACHING SUGGESTIONS (CONTINUED)

Name	Paraffin	Dextrose	Sucrose	Salt
Formula	$C_{25}H_{52}$	$C_6H_{12}O_6$	$C_{12}H_{22}O_{11}$	NaCl
Bond type	Covalent	Covalent	Covalent	Ionic
Polarity	Nonpolar	Polar	Polar	X
Electrons	Total: 202 $n = 1, 2$	Total: 96 $n = 1, 2$	Total: 182 $n = 1, 2$	

Students may need assistance with the formula for paraffin because it is usually considered a range of alkanes. In the materials for this lesson, the formula of  $C_{25}H_{52}$  is used for paraffin. As students work, walk around and monitor their progress with establishing the appropriate relationship between melting point and intermolecular forces and completing the table accurately. It is recommended that you display or project the table and establish a consensus before proceeding with the lesson. An example is included as Table A.

Question 3 is designed to push students one step closer to the concept of intermolecular forces. Use this opportunity to launch into a brief lesson on intermolecular forces. The following paragraphs address the essentials of the content, and there is a space in the student pages for note-taking to occur. The information in the lesson must adequately prepare students to answer the remaining Analysis and Conclusion Questions. Ideally, the lesson is more of a dialogue and less of a lecture. Examples of teacher questions are embedded in the following content.

Q. What is the difference between the charges held by ions and the dipoles found in polar molecules? Does it make sense that salt had the highest melting point of the four substances?

The key idea is that a dipole represents a “partial” charge as opposed to the full charge possessed by ions. These “partial” charges can also establish coulombic forces but at a lesser magnitude of force than the “full” charge of ions. Connecting this concept to the lab, salt has the greatest melting point because disrupting the electrostatic attractions between ions requires significantly more energy to mobilize those ions out of the solid structure.

Ionic compounds do not form molecules but rather are held together in a crystalline structure by omnidirectional electrostatic attractions. Intermolecular attractive forces are appropriately discussed when ions are in solution but not when ionic compounds are in the solid state.

## TEACHING SUGGESTIONS (CONTINUED)

Q. What does the prefix “inter-“ with the root “molecule” tell us about the type of forces described as “intermolecular forces”? Why does Coulomb’s law apply to these forces? How do these forces affect the way molecules behave around each other? What governs the strength of these forces?

For covalent substances, the forces that hold one molecule to another molecule are referred to as *intermolecular forces (IMF)*. These forces arise from the unequal distribution of the electrons in the molecule and the electrostatic, or *coulombic*, attraction between oppositely charged portions of neighboring molecules.

Some molecules are naturally polar, therefore they have permanent partial positive and partial negative portions of the molecule. The type of attractive force between oppositely charged portions of associating molecules is referred to as *dipole-dipole interactions*.

It should be noted that the portions of neighboring molecules with like partial charges repel each other, and it is the balance between the attractive and repulsive forces that governs the physical arrangement of a polar molecule in a system of other polar molecules. The magnitude of the net dipole, which is affected by the size of the difference between electronegativities and the distribution of bond dipoles, is proportional to the strength of the interaction.

Using a visual representation of a generic polar molecule, such as the one shown in Figure A, can help students grasp this concept. Note that this section of content contains the information pertinent to answering Question 4.

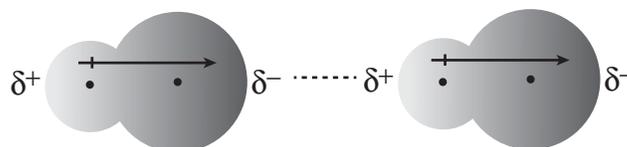


Figure A. The partial negative charge on the left molecule experiences an attractive force with the partial positive charge on a neighboring molecule.

Q. Which of the molecules investigated have dipole-dipole interactions? What makes them identifiable?

Direct students to Table 1 and instruct them to add a row labeled “Intermolecular Forces.” Students should identify which molecules exhibit dipole-dipole interactions. Note that sodium chloride is not a molecule and therefore does not qualify for consideration.

## TEACHING SUGGESTIONS (CONTINUED)

**Q.** If electronegativity is a significant factor in determining polarity, where are the elements with the greatest electronegativities on the periodic table? What happens when a hydrogen atom is covalently bonded to one of the three most electronegative elements? What conditions are necessary for hydrogen bonds to occur? How does hydrogen bonding compare to dipole-dipole interactions? What is the difference between a bonded hydrogen and a hydrogen bond?

*Hydrogen bonding* is a type of intermolecular attraction that is only available to a select subset of polar covalent molecules. 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. As 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 a lone pair of electrons on a highly electronegative atom (nitrogen, oxygen, or fluorine) in another molecule.

In the classic example of a water molecule, the hydrogen bonding is significant because there are two hydrogen atoms and two lone pairs on every molecule. This is not to be confused with the two bonded hydrogen atoms a water molecule has through the *intramolecular* sharing of valence electrons in the covalent bond with oxygen. Hydrogen bonding is often considered a more extreme version of dipole-dipole interactions because the mechanism of attraction is similar.

Comparative analogies can help students with this relationship. They can even make their own analogies. Examples include:

- Hydrogen bonding is to dipole-dipole as running is to jogging.
- Hydrogen bonding is to dipole-dipole as varsity is to junior varsity.
- Hydrogen bonding is to dipole-dipole as boiling is to evaporating.
- Hydrogen bonding is to dipole-dipole as a freezer is to a refrigerator.

Hydrogen bonding is significant in determining such factors as the high boiling point of water, the solubility of acetone in water, and the shape and structure of proteins and DNA. Note that the content of this section contains the information pertinent to answering Question 5.

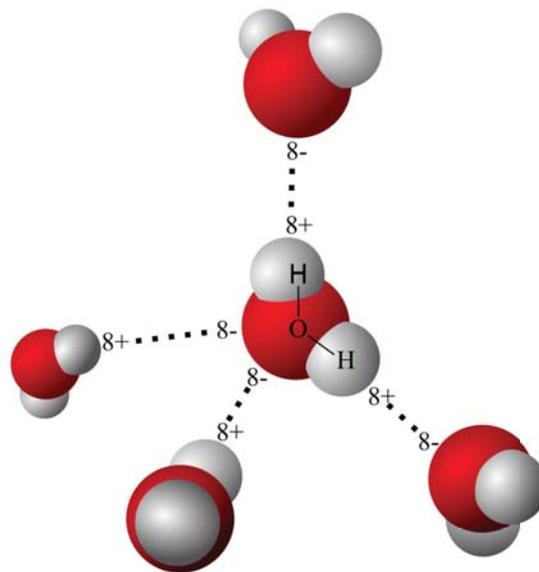


Figure B. The partial positive charge of each hydrogen atom is strongly attracted to a lone pair on the oxygen atom of a neighboring molecule, which is called hydrogen bonding.

## TEACHING SUGGESTIONS (CONTINUED)

Q. Which of the molecules you investigated exhibit hydrogen bonding? What makes them identifiable?

Direct students to Table 1 and ask them to determine whether any of the molecules would exhibit hydrogen bonding. Both dextrose and sucrose should be labeled as having hydrogen bonds.

Q. Particles within ionic compounds and molecules of polar covalent compounds interact with ways that are clearly described by Coulomb's law, but what about molecules with no net polarity? What is responsible for molecules of paraffin, a nonpolar substance, holding together in a solid state at all at room temperature?

All molecules experience a type of intermolecular force called *London dispersion forces*. These interactions 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 oxygen,  $O_2$ ). Pretend that the molecule is completely 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 many other molecules that are constantly colliding with it. When these collisions occur, the electron cloud around the molecule is distorted. This produces a momentary instantaneous dipole within the molecule. The extent of distortion of the electron cloud is referred to as *polarizability*.

Because the molecule now has a positive side and a negative side, it can be attracted to other molecules including neighboring nonpolar molecules that acquire an induced dipole as a result of their proximity to the instantaneous dipole. This attractive force is called a *London force* or a *dispersion force*. It can be described as the attractive force between temporary partial opposite dipoles on neighboring molecules. As all molecules have electrons, all molecules have London forces.

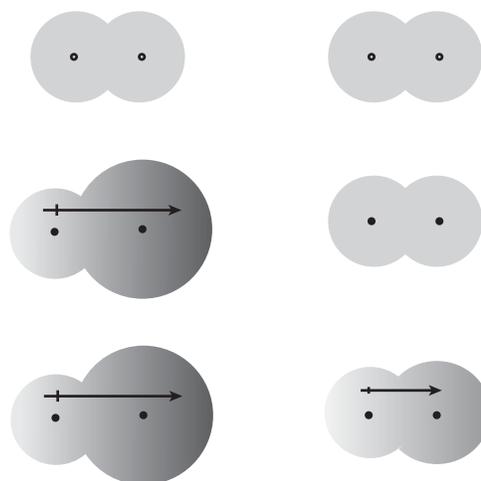


Figure C. Two nonpolar molecules. When the molecule on the left develops an instantaneous dipole, it induces a dipole in the molecule on the right.

**TEACHING SUGGESTIONS (CONTINUED)**

To illustrate how London dispersion forces alone can hold nonpolar particles together in a liquid or solid state, an analogous scenario can be presented to students. Imagine a completely still body of water. Now imagine a single raindrop hitting the surface of that water. At the point of contact, the water becomes instantaneously polarized, no longer evenly distributed. This effect ripples outward radially, representing the phenomenon of induced polarizability. Naturally this effect is temporary and after time, even as the ripple spreads, the surface at the origin of the polarizability will return to being still again.

Now imagine that instead of one drop hitting the still surface, there is a steady rain. There are now a great many sources of “instantaneous dipoles” from which “induced dipoles” can ripple. The surface of the water is never still because as an area attempts to reach stillness, it is quickly polarized again by another ripple. Although the polarization is temporary, it is happening with such frequency it is virtually constant. Therefore, depending on how polarizable the electron cloud is, the net forces can be strong enough to hold particles together in a solid state. Note that the content of this section contains the information pertinent to answering Question 6 and Question 7.

Q. Which of the molecules that you investigated exhibit London dispersion forces? What makes them identifiable?

Direct students to Table 1 and ask them to determine whether any of the molecules would have London dispersion forces. All three covalent compounds should be labeled as having London dispersion forces because they all have electrons.

Visual representations of the “sum of intermolecular forces” can help students with the idea of how the prevalence of a force is often more important than the nature of the force itself. Table B shows an example of a simple visual vector representing the different intermolecular forces present in a sample of molecules. The solid lines represent the net forces attributed to London dispersion forces.

Because all of the molecules being compared are organic molecules with electrons held in the first and second energy levels only, the number of electrons in the molecule determines the relative heights of the lines. Although the shape of the molecule also affects the polarizability, for the sake of simplicity the electron count is sufficient for modeling relative proportion. The dashed line represents the net forces attributed to the polarity of the molecule.

## TEACHING SUGGESTIONS (CONTINUED)

In the example provided in Table B, no distinction is made between dipole-dipole interactions and hydrogen bonding because it is not easily differentiated in terms of magnitude. Because both dextrose and sucrose have hydroxyl (–OH) groups

as well as double-bonded oxygen atoms, the number of hydroxyl groups can inform us that the net forces due to polarity and hydrogen bonding in sucrose should be relatively greater than those for dextrose.

Name	Paraffin	Dextrose	Sucrose	Salt
Formula	$C_{25}H_{52}$	$C_6H_{12}O_6$	$C_{12}H_{22}O_{11}$	NaCl
Bond type	Covalent	Covalent	Covalent	Ionic
Polarity	Nonpolar	Polar	Polar	
Electrons	Total: 202 $n = 1, 2$	Total: 96 $n = 1, 2$	Total: 182 $n = 1, 2$	
Intermolecular force	LDF	LDF Dipole-dipole Hydrogen bonding	LDF Dipole-dipole Hydrogen bonding	
Visual representation of magnitude of intermolecular interaction				

## TEACHING SUGGESTIONS (CONTINUED)

In some cases, the net contribution by London dispersion forces outweighs the attractions between permanent dipoles. In Conclusion Question 2, students are asked about water. Despite having hydrogen bonding water is a liquid at room temperature, indicating lower intermolecular

forces than even the nonpolar paraffin. The same exploration of visual vectors can help students see that the large number of electrons, and therefore greater polarizability of the electron cloud, makes the intermolecular attractions stronger in paraffin without permanent dipoles or hydrogen bonds.

Table C. Properties of Water vs. Paraffin		
Name	Water	Paraffin
Formula	H <sub>2</sub> O	C <sub>25</sub> H <sub>52</sub>
Bond type	Covalent	Covalent
Polarity	Polar	Nonpolar
Electrons	Total: 10 <i>n</i> = 1,2	Total: 202 <i>n</i> = 1,2
Intermolecular force	LDF Hydrogen bonding	LDF
Visual representation of magnitude of intermolecular interaction		

**ANSWER KEY****PRE-LAB EXERCISES**

1. Electrons and electrons: repulsive force.  
Electrons and protons: attractive force.  
Protons and protons: repulsive force, unless close enough for the strong nuclear force to hold them together.
2. A strong attraction results from the transfer of electrons from one atom to another. The less electronegative atom will give up its electron, and the more electronegative atom will gain it. An atom that has lost its electron will have a positive ionic charge whereas an atom that has gained an electron will have a negative ionic charge. As opposite charges attract, the ions that result from a transfer of an electron will be attracted to each other.
3. Energy is required to move two positively charged particles together because similar charges repel each other. According to Coulomb's law, the force is inversely related to the square of the distance. Therefore, the force of repulsion will increase as the distance between the particles decreases. To overcome this repulsive force and keep the particles close together, energy must be applied to the particles.
4. A "physical property" is one that can be measured without changing the chemical composition of the substance. Density and melting point are two examples of physical properties. Other examples include but are not limited to color, luster, malleability, boiling point, vapor pressure, volume, and mass.
5. During the melting of a solid, particles break free of a specific structural arrangement and are able to move more freely among other particles.

**DATA AND OBSERVATIONS**

Paraffin, dextrose, sucrose, and salt.

## ANSWER KEY (CONTINUED)

## ANALYSIS

- In general, stronger intermolecular forces should result in a higher melting point.
- See Table 1.

Name	Paraffin	Dextrose	Sucrose	Salt
Formula	$C_{25}H_{52}$	$C_6H_{12}O_6$	$C_{12}H_{22}O_{11}$	NaCl
Bond type	Covalent	Covalent	Covalent	Ionic
Polarity	Nonpolar	Polar	Polar	X
Electrons	Total: 202 $n = 1, 2$	Total: 96 $n = 1, 2$	Total: 182 $n = 1, 2$	
Intermolecular forces	LDF	LDF Dipole-dipole Hydrogen bonding	LDF Dipole-dipole Hydrogen bonding	

- Molecular attractions are not as strong as ionic attractions. According to Coulomb's law, the magnitude of the charges is directly proportional to the force of attraction.  
At the same distance, ions will have stronger attractions because they have absolute charges whereas the most that a molecule can achieve with a dipole is a partial charge.
- The key factor that affects the strength of attractions is the magnitude of the molecular dipole as determined by the difference in the electronegativity values between bonded atoms and the distribution of electron density across the molecule.
- Hydrogen bonding is a more extreme version of a dipole-dipole interaction. Both forces rely on differences in electronegativities between atoms in polar molecules, but hydrogen bonding is established only when there is a hydrogen bonded to nitrogen, oxygen, or fluorine on one molecule and a lone pair of electrons on nitrogen, oxygen, or fluorine of a neighboring molecule.  
Dipole-dipole interactions will occur any time two polar molecules are neighboring, whether or not they have hydrogen, nitrogen, oxygen, or fluorine.

**ANSWER KEY (CONTINUED)**

6. Molecular collisions cause electron density to shift in a molecule, creating an instantaneous dipole, or polarity. This momentary polarity causes induced dipoles in neighboring molecules.

The temporary positive dipole on one molecule is attracted to the temporary negative dipole on a neighboring molecule. The strength of the attractive force depends on how easily and by how much the electron cloud can be distorted. Having a greater number of electrons and having electrons further from the nucleus increases the polarizability of the electron cloud.

As all molecules have electrons, all molecules experience London dispersion forces.

7. Polarity is a measure of the uneven distribution of electron density in a molecule. Polarizability is a measure of the ability to affect the distribution of electron density in a molecule. Nonpolar molecules can be made temporarily polar if the molecule is polarizable.

**CONCLUSION QUESTIONS**

1. Of the four compounds, paraffin has the lowest melting point followed by dextrose and sucrose, respectively.

Salt had the highest melting point by default because it never melted. During melting, particles must overcome attractions to other particles. Because the ions in salt have absolute charges, the strength of that attraction is very high and requires the greatest amount of energy to overcome.

Paraffin is a nonpolar molecule. It has more than twice the number of electrons as dextrose, and therefore more London dispersion forces. However, dextrose is a polar molecule that also experiences hydrogen bonding, which is why the total intermolecular attractions outweigh the London dispersion forces experienced by paraffin.

Like dextrose, sucrose is polar with hydrogen bonding but sucrose has almost twice as many electrons as dextrose, which makes it more polarizable. Sucrose also has more hydroxyl (–OH) sites available for hydrogen bonding. These two factors make the intermolecular attractions greater than dextrose, causing sucrose to require more energy to overcome the attractions. The more energy required to overcome attractions, the higher the melting point.

**ANSWER KEY (CONTINUED)**

2. Paraffin is a nonpolar molecule, and the only intermolecular force between its molecules is the London dispersion force. Water is a polar molecule and experiences both London dispersion forces and hydrogen bonding.

Paraffin has 202 electrons per molecule whereas water has 10 electrons per molecule. Both molecules have electrons in the first two energy levels. The strength of the London dispersion forces depends on the number of electrons. Therefore, the number of electrons in paraffin contributes to a magnitude of force that outweighs the total intermolecular forces in water, even with the presence of hydrogen bonding.

3. Both water and hydrogen sulfide molecules are polar with a bent molecular geometry. Hydrogen sulfide has 18 electrons that can be found in the first three energy levels whereas water has 10 electrons that can be found in the first two energy levels. Based on the number of electrons and the size of the electron cloud, the polarizability and thereby the strength of the London dispersion forces of hydrogen sulfide is greater than that for water.

Water molecules experience hydrogen bonding whereas hydrogen sulfide molecules experience dipole-dipole interactions. The contribution from the two hydrogen atoms engaging in hydrogen bonding make the attraction between water molecules greater than that for hydrogen sulfide, despite the fact that water has weaker London dispersion forces.

4. All of these diatomic halogens are nonpolar molecules. As you move down the column of halogens in the periodic table from fluorine to iodine, the number of total electrons and the size of the electron cloud increases, which increases the polarizability of the electron cloud. A more polarizable electron cloud results in stronger London dispersion forces between molecules.

Fluorine and chlorine have the weakest intermolecular forces, so these are gases.

Bromine has stronger forces that hold molecules together in the liquid state.

Iodine, having the strongest intermolecular attractions from London dispersion forces, has molecules held together in a solid state at room temperature.



NATIONAL  
MATH + SCIENCE  
INITIATIVE

#### MATERIALS

aprons

goggles

hot plate

paper towels

can lid

marker, Sharpie®

dextrose

paraffin wax

sucrose

wood splint

salt

# Don't Flip Your Lid

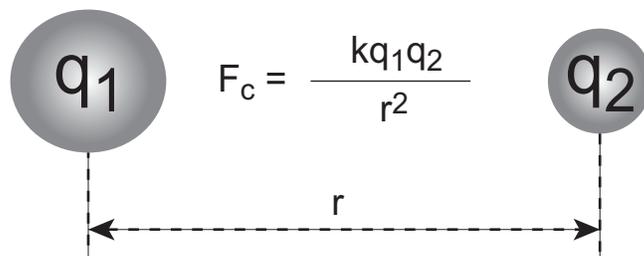
## Comparing Intermolecular Forces

The interaction of particles governs the physical world. The competing forces of attraction and repulsion between subatomic particles are responsible for properties of the atom that include size, ionization energy, and electronegativity.

Between two atoms, those same forces influence the way valence electrons associate with each atom. Electrons can be transferred or shared to create an alliance, or **bond**, between two atoms. When electrons are transferred, the resulting ions are so strongly attracted to each other that they arrange themselves into a crystalline structure that maximizes those attractions. These interactions can all be understood qualitatively by **Coulomb's law**: The force of the attraction, or repulsion, is related to the magnitude of the charges,  $q_1$  and  $q_2$ , and the distance between the particles (Figure 1).

The same principle that holds cations and anions together also gives us a way to understand the physical properties of substances. Metallurgy, cooking, the development of plastics and polymers, petroleum engineering, and even candle making are examples of where differences in the physical properties of the materials used and created have application.

Figure 1. Coulomb's law



In physical processes like phase changes, collections of particles interact without any change in chemical composition. The nature of these interactions depends on a set of factors that influence the degree of coulombic attractions between particles. Understanding these factors allows us to make informed predictions about the relative properties of pure substances, including the melting point.

## PURPOSE

In this laboratory activity, you will determine the relative melting points of four substances: paraffin, sodium chloride (table salt),  $C_6H_{12}O_6$  (glucose), and  $C_{12}H_{22}O_{11}$  (sucrose).

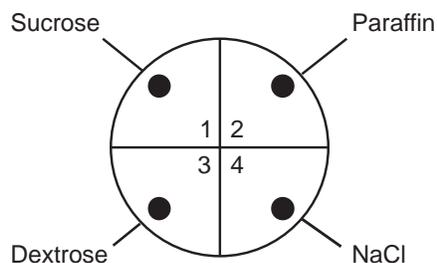
### SAFETY ALERT!

- » Wear your goggles.
- » Allow the can lid to cool before disposal.

**PROCEDURE**

1. In the space marked "Hypothesis" on your student answer page, write a statement predicting the order in which the salt, paraffin, sucrose, 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.

Figure 2. Can lid with four substances



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.
4. Carefully place the can lid onto an unplugged hot plate. Plug in the hot plate and turn the heat up to a medium setting.
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 hot plate and unplug it.
6. Allow the can lid to cool and dispose of it in the trash.



**HYPOTHESIS****DATA AND OBSERVATIONS**

Record the relative melting for the four compounds in order of increasing melting point.

**ANALYSIS**

1. An *intermolecular force* is defined as the force of attraction between neighboring molecules. Considering this definition and what you know about melting points, what is the general relationship that exists between the strength of intermolecular forces and melting point?
  
2. In Table 1, write the names and formulas of the four substances from left to right in order of increasing strength of intermolecular forces. Identify the bond type as either covalent or ionic. Finally, complete the rest of the table for covalent molecules only.

<b>Name</b>				
<b>Formula</b>				
<b>Bond type</b>				
<b>Polarity</b>				
<b>Electrons</b>				
<b>Intermolecular forces</b>				

**ANALYSIS (CONTINUED)**

3. Coulomb's law describes why oppositely charged particles are attracted to each other. Using the concept of dipoles, explain why the attractive forces between neighboring molecules are not as strong as the attractive forces between ions.

**NOTES ABOUT INTERMOLECULAR FORCES**



**CONCLUSION QUESTIONS**

1. From what you know about attractive forces between particles, explain the relative order of melting points for the four substances investigated.
2. In determining the strength of the intermolecular forces, it is important to consider the net forces present. Use this argument to defend why water, despite having hydrogen bonding, has a much lower melting point than paraffin.
3. Hydrogen sulfide is a gas at room temperature whereas water is a liquid, yet hydrogen sulfide has more electrons than water. Explain this anomaly.
4. Consider the halogens at room temperature and 1 atmosphere of pressure. Explain why fluorine and chlorine are gases at room temperature whereas bromine is a liquid and iodine is a solid.



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