Chemistry - Unit 6  Sticky Tape Activity

Part 1 – Preparing the tapes - examining their behavior
Place a 15 cm piece of transparent tape on the table; this base tape serves to keep the bottom tape clean and remains on the table. Two more pieces of tape are placed on the base tape (see figure below). In the same way, prepare a second set of tapes.
The top-bottom pair is removed slowly from the base tape, then grounded by gently rubbing one’s finger down the length of the top tape. Once each pair of tapes appears to behave like an uncharged piece of tape, the top and bottom tapes are separated quickly.

Attach the strips (foil, paper, top and bottom) to the edge of the table; this will enable you to bring the other tapes and strips near in order to test for the presence and magnitude of electrical interactions between all combinations. See the figure below.

Record your findings in the table below.

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Part 2 - The assignment of (+) and (-) charges

After you have summarized your findings (attraction, repulsion or no interaction) for the two tapes, foil and paper, rub a hard rubber or plastic rod with fur or wool. Approach the T tape, the B tape, the foil and the paper with the rod. Describe what you see. Does the rod behave more like the T tape or the B tape?

<table>
<thead>
<tr>
<th>Plastic/wool</th>
<th>Top</th>
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<th>Foil</th>
<th>Paper</th>
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Based on a number of observations scientists have assigned the label of negative (–) to the charge of a rubber or plastic rod rubbed with fur or wool. The fur or wool becomes positively charged (+). Based on your observations from using the rod, label the T and B tapes as either a (+) or (–).

Part 3 - the Thomson model of the atom

Now, reflect on the relative strengths of the interactions between the tapes, foil and paper. Rank the strength of attraction between the T tape and the B tape, T tape and foil and T tape and paper. Your instructor will help you develop model of the atom that helps to account for these interactions.
Chemistry – Unit 6 Notes
Thomson Model of the Atom

J. J. Thomson performed experiments with cathode rays in an attempt to understand electricity – which was still a mystery in the late 1800s. Review the website A Look Inside the Atom¹ to find the conclusions that Thomson and other physicists drew regarding the mysterious cathode rays.

Thomson’s 1897 Experiments - state the conclusions Thomson drew from each of his famous cathode ray experiments:

1. **First Experiment:** Thomson directed the beam at an electrometer and tried to separate the evidence of charge from the path of the beam. *What connection did Thomson find between charge and the cathode rays? Was the charge positive or negative?*

2. **Second Experiment:** Thomson tried passing the cathode ray through an electric field. *How did cathode ray beam behave when it passed through an electric field? What did he conclude after his second experiment?*

3. **Third Experiment:** Thomson did some careful measurements on how much the path of the cathode ray was bent in a magnetic field and how much energy they carried. From this work Thomson could describe the mass/charge ratio of the cathode ray particles. *What amazing result did Thomson find?*

**Thomson’s Atomic Model:** Thomson presented three hypotheses from his experiments. Only two were accepted by physicists – in fact the third was shown to be wrong! From the first two came a model of the atom known as the *Plum Pudding* model. Complete the atom drawing below to illustrate Thomson’s plum pudding model. Explain how this fits with his observations.

¹ http://www.aip.org/history/electron/jjhome.htm
Let’s see how we can use Thomson’s model to explain the behavior of the sticky tape when we made our tape stacks.

A few atoms from the top tape and the bottom tape are represented in the diagram below. Add electrons to each atom to show what happens to the electrons when we make a tape stack out of neutral pieces of tape and then pull them apart.

Describe the macroscopic changes in the tapes and then provide a microscopic explanation based on Thomson’s model of the atom and your drawings.

Behavior of Foil and Paper with Charged Tapes

We observed that neither foil (metal atoms) nor paper (non-metal atoms) would attract each other. But foil and paper are both attracted to both the charged tapes (top and bottom).

*How can we use the pudding model of atoms to explain the differences we observed?*

Several atoms from the paper and foil are drawn on the next page. The ones on the left have no charged object near them. The ones on the right are next to a top tape (+ charge).
Add force vectors to the non-metal (paper) atoms and the top tape in the first row to show the attraction between the paper and the tape. Then do the same for the foil and the tape in the second row. Be sure the size of the vectors shows the relative strengths of the attractions.

Now draw the electrons in each atom “bowl” to show their arrangements when no charged object is near present and then when a charged object is brought near.

Explain why these arrangements of electrons would produce the observed attractions.
Unit 6 – Particles with Internal Structure
The Role of Charge and Nomenclature

Instructional goals

1. Describe the evidence that supports the idea that the simple particles have a property we call charge.

2. Describe the evidence that led Thomson to suggest that the mobile charge in atoms is negative.

3. Use the Thomson model of the atom to account for the fact that neutral atoms can become either positively or negatively charged by the loss or gain of electrons.

4. List properties that distinguish metals from non-metals.

5. Describe the evidence that distinguishes ionic from molecular or atomic solids.

6. Given the formula of an ionic or molecular substance, state its name.

7. Given the name of ionic or molecular substance, write its formula.

8. From the name or formula of a substance determine whether that substance is ionic or molecular.

Sequence

1. Sticky tape lab

2. Post-lab discussion: Thomson’s experiments and model of the atom

3. Application of Thomson model to conductivity and polarization

4. Activity: conductivity of solutions

5. Electrolysis of CuCl₂

6. Patterns of Charge in the Periodic Table - Worksheet 1

7. Structures of solids – activity using Mercury software and crystal structure data

8. Implications of structure – notes and worksheet 2

9. Ionic nomenclature - Worksheet 3

10. Molecular nomenclature – Worksheet 4

11. Quiz

12. Worksheet 5 – Representing empirical formulas
13. Worksheet 6 (optional) – more practice with empirical formulas

14. Begin nail lab – Unit 6 review

15. Part 2 of nail lab - finish unit review

16. Unit 6 test

Overview

In Unit 4 our model of the atom moved from simple BB’s to one in which some substances were made from “compound particles”. Using the electrolysis of water, we showed that these particles combined in definite ratios. What we didn’t address, however, was what held these particles together in these well-defined ratios.

The phenomena we’ve studied thus far did not require that these particles have any internal structure. However, in the electrolysis of water, it is clear that electrical forces are somehow involved in the formation of compounds. In the Sticky Tape lab we study the behavior of the charged particles and develop a more complex model of the atom that accounts for the fact that some particles have positive charge whereas others are negatively charged.

Most high school chemistry texts delve into an in-depth treatment of the internal structure of the atom very early on. Perhaps the authors feel that students need to accept the model of the nuclear atom in order to learn how to write formulas of ionic compounds. We believe this approach to be unwise because it requires students to accept a more complicated model of the atom than is necessary to account for phenomena they can observe. If students know that atoms are composed of smaller fundamental particles, it is because their earlier science teachers and texts have simply asserted that this is the case. Few, if any, students at this stage can cite any evidence for their belief in electrons. The NSES Content Standard B (Physical Science) urges us not to offer models without any of the evidence that lead to their development.

"It is logical for students to begin asking about the internal structure of atoms, and it will be difficult, but important for them to know "how we know." Quality learning and the spirit and practice of scientific inquiry are lost when the evidence and argument for atomic structure are replaced by direct assertions by the teacher and text. Although many experiments are difficult to replicate in school, students can read some of the actual reports and examine the chain of evidence that led to the development of the current concept of the atom." (p 177)

The Sticky Tape lab provides the evidence to support a model of the atom with a positive core and negatively charged particles that are mobile to varying degrees in various materials. This model helps us to account for the differences in the electrical conductivity of metals and non-metals. Later, it helps us to distinguish ionic solids from molecular solids. In the former, the fundamental structure is a lattice of oppositely charged ions held together by strong electrostatic forces. These account for the relatively high melting and boiling points of these substances. When ionic solids are melted, the melt conducts electricity because the charged particles, previously tightly bound in the solid, are now free to move about. By contrast, molecular solids are composed of discrete molecules held together by relatively weak dispersion forces or by stronger dipole-dipole forces. When a molecular solid is melted, the basic structural units - electrically neutral molecules - do not conduct electricity.

Finally, students learn general rules of nomenclature, relating names to the formulas of either ionic or molecular compounds. The difference in the rules reflects the different types of
compound formed by the atoms. Students are given opportunities to draw particle models of ionic compounds and the ions from which they are formed so that they can relate symbolic with diagrammatic representations of structure.

Instructional Notes

1. Sticky Tape Lab

Apparatus
For each 2 person group:
One roll of transparent tape (recommended: 3M Magic tape, may be shared between 2 groups)
Two strips of paper, 15 cm x 1.5 cm
Two strips of aluminum foil, 15 cm x 1.5 cm
Plastic ruler, glass rod
10 cm x 10 cm pieces of wool or fur, cloth

Pre-lab discussion
- Remind the students that there was something about water than caused the constituent gases to collect at different charged electrodes. Perhaps electrical charge is somehow involved in the way the hydrogen and oxygen are held together in water. You can demonstrate the existence of the electrical force by rubbing a balloon against your hair and then allowing it to stick to a wall. At this point, it might be worthwhile to investigate the nature of this electrical force.
- Demonstrate that pulling two pieces of tape apart (see procedure notes 1-5) produces tapes that can interact with an electrical force. Guide students in a discussion of how to investigate interactions with the other materials supplied.
- Students should work in pairs with a roll of tape for each group. After showing how to create usable strips of tape, the students, in the true modeling way, should be challenged to explore all possible two-strip interactions as well as interactions between the strips and other student-chosen objects (such as pens, notebooks, themselves, etc.).
- Emphasize that observations must include a written description and include a series of sketches of the tapes as they approach one another with vectors to represent the forces on the tapes. Label the forces. A possible data table appears below.

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Performance notes

Two pieces of tape are placed on the base tape (see figure below). The top-bottom pair is removed *slowly* from the base tape, then grounded by gently rubbing one’s finger down the length of the top tape. Once each pair of tapes appears to behave like an uncharged piece of tape, the top and bottom tapes are separated *quickly*. The base tape, which serves to keep the bottom tape clean, remains on the table.

![Diagram of tape setup](image)

The strips (foil, paper, top and bottom) should be attached to the edge of the table to enable the students to bring the other tapes and strips near in order to test for the presence and magnitude of electrical interactions between all combinations.

**NOTE:** The neutral paper and foil are attracted to both top and bottom tapes. Neutral does not mean *no* charge; in fact, the neutral paper has billions of charges, it’s just that the + and – charges are approximately the same in number and evenly distributed throughout the object so they neutralize each other. The proximity of a charged object has the effect of re-distributing the charges within the neutral object. This effect is called polarization. The effect is somewhat different for conductors and insulators. In conductors, electrons actually relocate from one side of the object to the other. In insulators the electrons, while bound to specific nuclei, spend more time on one side of the nuclei than on the other. You will have the opportunity in the post-lab discussion to help students build a model of the internal structure of the atom that accounts for this phenomenon. Note that when two objects are electrically attracted to each other, this does NOT confirm that both objects have a NET charge on them.

2. Post-lab discussion – need for a more detailed model of the atom

1. **Preparation:** Look for reasonable data tables describing the interactions during group presentations. Students should clearly identify the two objects causing the interactions noted. Try to reach consensus about the kinds of interactions observed. Have students repeat trials if necessary.

2. **Summarize observations:** Note that all the objects show one of two types of behavior:
   a) three attractions (one strong) and one repulsion
   b) two attractions and two no effects

3. **Construct a descriptive macroscopic model:** Looking at the above behaviors, it is clear that the repulsive interaction is a unique identifier for the type of behavior an object will exhibit in all
other interactions. If an object repels either T or B tapes, it will attract all the three other strips; if it is attracted by both tapes, it will not interact with the paper or foil. To model this behavior, we assign a property we call charge to an object capable of exerting a repulsive force. Since B tapes repel one another (as do T tapes), we propose that like charges repel and opposite charges attract. All that remains is to decide the sign of the charges on each of the tapes. Instruct the students that scientists have assigned a (−) charge to plastic rubbed with fur or wool\(^1\). Have them go back and test the interactions of these charged objects with the tapes, foil and paper. They should find that the (−) plastic repels the B tape and attracts the other three strips; hence behaving like the B tape. Let the students add these interactions to their table. They should find that these definitions form a coherent pattern in all the data rows: opposite charges attract; like charges repel; neutral objects are attracted to both charges, but exert no force on other neutral objects.

4. **Construct an explanatory microscopic model:** To explain the source of these charges, we need to expand our model of the atom to have some internal structure. We will assume that each atom contains both positive and negative charges that normally cancel each other. Introduce the evidence that lead J. J. Thomson to propose that in solids, only the negative charges are free to move, and that these charges are much smaller than an atom and carry only a negligible fraction of its mass. The “[A Look Inside the Atom]\(^2\)” website traces some of Thomson’s experiments with cathode rays. The video **Smaller Than the Smallest**\(^3\) is another resources that shows how the results of JJ Thomson’s experiments with cathode rays brought him to the realization that the atom was no longer the smallest constituent of matter. His plum-pudding\(^4\) model accounted for much of the electrical behavior of atoms. We will use the Thomson model of the atom – a massive positive core associated with a small number of mobile, negatively charged particles we call “electrons”. A visual representation of this model is the “plum pudding” – the positive cores are represented by bowls of pudding, which attract the negative electrons represented by raisins. The attraction of the pudding for raisins in some of the bowls is stronger than in others; raisins can move from one bowl to another because of such differences in attraction. However, since raisins also repel one another, you cannot cram too many raisins into the same bowl of pudding. A study guide “01a_Thomson model notes” is provided in the Unit 6 folder to help students organize what they learn from the American Institute of Physics website.

5. **Application of the Thomson model to the separation of sticky tapes:** To begin a discussion of the concept of conservation of charge, have the students consider more carefully what occurs when the top and bottom tapes are pulled apart. Based on our findings, the figure below might represent a layer each of atoms in the top and bottom tapes:

![Diagram of atom layers](image)

When two objects of different substances come into contact, some electrons move from one substance to the other, based on the relative attractions of the cores of the two substances (some

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\(^1\) An optional trial is to test the strips with glass rubbed by plastic (+); it repels the T tape and attracts the other three.

\(^2\) [http://www.aip.org/history/electron/jjhome.htm](http://www.aip.org/history/electron/jjhome.htm)


\(^4\) Since students are unlikely to know that plum-pudding is more like a fruitcake, we’ll use raisins to represent the mobile negatively charged particle.
raisins creep from one set of bowls into the other). If the objects are then quickly separated (rubbing is just a repetitive contact and separation\(^5\)), an excess of electrons remains in one object, counterbalanced by a deficiency of electrons in the other (one set of bowls is now “raisin rich”, while the other is “raisin poor”). This microscopic imbalance of charges translates to an overall macroscopic charge on the object. The T tape becomes positively charged because electrons are transferred to the B tape. The overall number of electrons does not change, just their distribution on the tapes. Neutral atoms have the same amount of (+) and (−) charge. Be aware that your students may have the naïve conception that electrons should be transferred to (not from) the top tape because they expect the adhesive on the top tape to pull electrons off the bottom tape. This is a good time to discuss the difference between what is going on at the macroscopic level (the adhesive is sticky to the touch) and at the microscopic level (the bowls on the dull side have a stronger attraction for electrons than do the bowls on the sticky side).

Emphasize that charge is not a substance, but a property of particles (cores and electrons) that determines the strength of their electrical interactions. It serves the same purpose as mass does in gravitational interactions. A study guide “01b_Sticky Tape notes” is provided in the Unit 6 folder to help students organize what they learn from the class discussion in this and the next section.

### 3. Further application of the Thomson model

#### Electrical conductivity

Begin the discussion with a demonstration of the difference in electrical conductivity of metals and non-metals using a simple electrical circuit with battery, leads and a bulb. An alternative is to use a portion of the video *Chemical Families* in which the conductivity of a number of elements is tested. In metals, electrons can be compelled to move from one core to another by the application of an external electric field provided by the battery. The same electric field does not result in movement of electrons in non-metals. We conclude that the attraction of the cores to the electrons is weaker in metals. This difference in behavior can be modeled by describing the pudding in metal as “soupy” and that in non-metals as “sticky”. Actually, the low energy required to move an electron from one metallic core to another serves as the basis to a more elaborate model of the bonding between metal atoms (quantum mechanical band theory), which regards the electrons as being delocalized across the entire metallic lattice.

#### Interaction between charged and neutral objects:

The most difficult aspect of the Sticky Tape lab is to have the students explain how a charged object can attract a neutral one. The distinction between conductors and insulators at the atomic level is needed to account for the fact that the attraction between both T and B tapes and the aluminum foil is greater than that observed with the paper. Remind the students that electrons can move readily from atom to atom in metals.

In a neutral foil, the electrons are equally distributed between atoms, and homogenously dispersed within each atom.

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\(^5\) http://www.electricityforum.com/static-electricity.html
When the positively charged tape is brought near the foil, electrons tend to move toward the side of the foil closest to the tape, making that side relatively negative ("raisin rich"). We say that the positively charged tape has polarized the foil because the electrons are no longer evenly distributed. The attraction between the (+) tape and the (-) side of the foil is stronger than the repulsion between the tape and the (+) side of the foil simply because the distance is smaller.

Have the students sketch the charge distribution that results when the (-) B tape is brought near the foil.

The effect of polarization in an insulator is less pronounced because the electrons are not so free to move about (each raisin stays in its own bowl). The (+) T tape can produce a shift in the electron distribution so that they are no longer symmetrically arranged about the core (the raisins are shifted towards one side of the bowl). The sides of the atoms in the paper nearest the tape become slightly negative charged, so a small attraction occurs.

Sample whiteboards are found in the miscellaneous folder. Fig 1 shows a good representation of the foil, but has the “raisins” moving from atom to atom in the paper. Fig 2 is a better representation. Fig 3 shows what a group thought of as a spark jumping the gap between foil and the tape.

The PhET website has a wonderful set of Java simulations to help illustrate some concepts that are ordinarily difficult to represent by static diagrams. The site also provides directions for how you can configure your computers to make use of them. One simulation: Balloons and Static Electricity (ballons_en.jar), does a great job representing the transfer of charge by contact and polarization of neutral materials. Another simulation: John Travoltage (travoltage_en.jar) helps to illustrate the mechanism by which one receives an electric shock when touching a conductor after picking up a surplus of electrons. These simulations are provided in the materials for this unit, but you are advised to visit the site to check out the other simulations available.

How do we apply this model to compounds? We can propose that electrical forces are involved in holding together the particles that make up pure substances. Perhaps the mobile negative charge is freer to move in some substances than in others. In the next activity students will find examine the electrical conductivity of solutions to provide evidence for this hypothesis.

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6 http://phet.colorado.edu/new/index.php
4. Conductivity of substances and solutions

**Apparatus**
Conductivity probes
Various substances (metals, molecular and ionic solids, liquids such as ethanol and vinegar)
100 mL beakers and distilled water

**Pre-activity notes**
Now that we have seen that various substance exhibit differences in the mobility of negative charge in the atoms, perhaps we can use this to sort compounds. Students can see for themselves that metals are conductors, but that the solids of compounds formed from metals do not conduct electricity. They will next examine the electrical conductivity of solutions of these compounds to determine whether the solutions exhibit the presence of charged particles. We can go on to further modify our atomic model of matter to account for our observations.

**Performance notes**
Have students test samples of metals and compounds for electrical conductivity. They should test the dry solids (where applicable) and then the solid mixed with water to produce about 25 mL of each solution in 100 mL beakers. The conductivity tests can also be carried out using spot plates if the conductivity electrodes are sufficiently close together to fit into the test well. Test the electrical conductivity of each solid and solution and record their findings; if possible, they should note relative conductivity of the solutions. If they test oil, they should do it last and wash off the electrodes thoroughly.

**Post-lab activity discussion**
Use conductivity to sort the compounds into electrolytes and non-electrolytes. Depending on the sensitivity of the apparatus, you may be able to discern strong from weak electrolytes. We observe two types of behavior among compounds. One type of substances does NOT conduct electricity either as a solid or as a liquid. We model these substances as being composed of neutral aggregates of atoms we call molecules. This behavior is typical of compounds of non-metals, but there are exceptions. The other type of substance does not conduct as a solid, but does when dissolved. We can model these as dissolving into oppositely charge particles we call ions that can move freely in solution. The presence of an electric field causes these charged particles to migrate resulting in an electrical current. We will examine this migration in greater detail in the next activity.

5. Electrolysis of copper(II) chloride

**Apparatus**
U-tube
2 carbon rods
10-25 mL 0.2 M CuCl2 (3.4 g CuCl2 • 2H2O/100mL)
9-V battery
leads to connect electrode to battery
conductivity probe
100 and 250 mL beaker

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7 These can be simply made from 9V batteries and LED’s, or purchased from vendors such as Flinn Scientific, or you can use conductivity probes by Vernier connected to the LabPro interface and computers or to a LabQuest.
Pre-activity notes
We have established that solutions of compounds formed by a metal and a nonmetal conduct electricity. We account for this conductivity in solution by proposing the existence of mobile charged particles which we call ions. Using the Thomson model of the atom, we can suggest that atoms that have lost one or more electrons become positively charged and those that have gained one or more electrons become negatively charged. During the electrolysis of copper(II) chloride we will examine the behavior of these aqueous ions to determine which are positive and which are negative. We use the Thomson model to explain our observation of what occurs at each electrode.

Performance notes
It is best if this activity is started at the beginning of the teaching period so that observations can be made both at the beginning and the end of class. First, demonstrate that solid copper(II) chloride does not exhibit conductivity, nor does distilled water. Test the copper(II) chloride solution for conductivity and ask students to explain this on a particle level based upon their previous experience with conductivity. Ask students to note the intensity of the color of the copper(II) chloride solution before they begin electrolysis; they will compare this to the color intensity afterwards. Within 10 minutes, student can see bubbles of a gas form at the anode; sufficient solid forms at the cathode for students to be able to tell that it is copper metal. Have students leave the apparatus running while they work on the discussion questions. They can return to their apparatus periodically to note further changes.

If possible, leave one apparatus connected overnight so that the students can observe an obvious loss of intensity in color of the solution and the accumulation of copper at the cathode. The changes are quite evident and they help to provide the basis for explaining what occurs in the Nail Lab. The quantity of chlorine gas produced is small, but to avoid possible irritation from the gas you may wish to have the students perform this experiment in a fume hood. Allow a few milliliters of the copper(II) chloride solution in an evaporating dish to evaporate overnight to illustrate that the solid copper(II) chloride will reform in the absence of water.

Post-lab activity discussion
Students should note that the copper metal accumulates at the negatively charged electrode; ask them to conclude what charge the copper ion must have. In like manner, the fact that chlorine gas accumulates at the positively charged electrode allows them to determine the charge on the chloride ions. Connect these observations to the fact that we explained conductivity by the migration of charged particles in solution. It is important for students to use our particle model to account for what takes place at each of the electrodes. Pose the following questions and ask students to try to sketch particle diagrams that help them answer the questions.

Why do the metal ions change from being soluble in solution to insoluble at the surface of the cathode?

Why do the nonmetal ions change from being soluble in solution to insoluble at the surface of the anode?

We want students to be able to explain that the (+) copper ions become neutral atoms when they pick up electrons at the cathode. At the anode, (–) chloride ions give up electrons to become neutral atoms that form diatomic molecules of chlorine gas.

8 Directions for how to download a short movie (96MB), describing how to perform this experiment, can be found on the Chemistry Core Units page at the AMTA website.
6. Patterns of charge in the Periodic Table – worksheet 1
From the combining ratios of the elements in the given compounds, students should be able to deduce the charge of ions in columns 1A, 2A, 3A, 6A and 7A. Try to get the students to reflect on the behavior of the top and bottom tapes when they describe the formation of cations and anions.

7. Structures of Solids – Mercury software activity
We have found that students have difficulty distinguishing between a formula unit, e.g., CaCO₃, that describes the empirical formula of an ionic solid, and a molecular formula, e.g., NH₃, that describes discrete molecules. In the September 2009 issue of *The Science Teacher*, Dennis Smithenry describes an activity that can help students better understand this distinction. It requires the use of crystal structure visualization software - Mercury – which you can download from the [Mercury](http://www.ccdc.cam.ac.uk/free_services/mercury/downloads) website. In the U6 resources folder you can find a copy of the article, step-by-step instructions on how to download and use Mercury, a zipped folder of the Teaching CIF files needed to use this resource, and an image from a screen capture for calcium carbonate.

You can use this software and images to help students distinguish between atomic, molecular, and ionic solids. Examples of the latter clearly show that the atoms in a formula unit are bound to surrounding atoms as well, whereas molecular solids have distinct molecules as their basic unit of structure. One way to use this resource is to run off the nine images on p55 of the article, laminate them and give them to groups of students. Then, you can display the structures of these nine substances, rotating them and allowing the students to see features of their structures. A better option would be to have the software and crystal structure files on the computers in the classroom so that students could examine these structures. Then give them the task of coming up with the rules for telling these classes of substances apart. Here's a quote from the article:

In a whole class discussion, students share their ideas and usually end up with some form of the following three rules:

1. Formula units contain more than one element and are bonded to other formula units.
2. Molecules contain more than one bonded atom and are not bonded to other molecules, except by weaker intermolecular forces.
3. Atoms of a single element may or may not be bonded to other atoms of that element.

Once they understand the rules, you can give them a few unknowns to practice on. A worksheet is provided to help students organize their observations.

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10 [http://www.ccdc.cam.ac.uk/free_services/mercury/downloads](http://www.ccdc.cam.ac.uk/free_services/mercury/downloads)
8. Implications of the structure of compounds – worksheet 2

The Mercury activity helps students to see that ionic and molecular compounds have different structures. Now it’s time to point out the implications of these differences.

Ionic compounds conduct electricity in the melt (and in aqueous solution), but not in the solid. We model these substances as lattices of positive ions (cations) and negative ions (anions) held together by electric attraction. In the solid phase, the mutual attraction of the ions keeps them locked in their lattice structure, and the application of an external electric field cannot dislodge them. However, in the liquid phase, the ions are free to move. The application of an electric field will cause the ions to move in a direction which depends on their charge. Cations, (+) ions are named because they are attracted to the (-) electrode (cathode) in a circuit. Anions, (–) ions derive their name from the fact that they are attracted to the (+) electrode (anode) in a circuit. This model fits other properties of ionic materials, such as their brittleness, high melting temperatures, and highly symmetric x-ray diffraction crystal structures. Since we already concluded that the attraction of electrons to the core in a metal atom is weaker than their attraction to non-metal cores, we expect the combination of metals and non-metals to form such ion pairs, in which some raisins from the metal pudding bowls transfer altogether into the non-metal pudding bowls (compare this behavior with the sharing of raisins between non-metal pudding bowls, where the attraction to both bowls is strong). Indeed, most compounds of metals and non-metals are ionic (but there are exceptions).

Ionic solids form crystal lattices according to the following rules
1. Ions are assumed to be charged, incompressible spheres.
2. Ions try to surround themselves with as many ions of opposite charge as closely as possible. Usually in the packing arrangement, the cation is just large enough to allow the anions to surround it without touching one another.
3. The cation to anion ratio must reflect the formula of the compound. In NaCl the lattice must be an array of chloride anions with the same number of sodium ions. In MgCl₂ the lattice contains only half that number of magnesium ions.

Each Na⁺ ion is equally attracted to the 6 Cl⁻ ions surrounding it. Remind students that at room temperature discrete molecules of NaCl do not exist. The ionic formula represents the simplest integer ratio (formula unit) of the cations and anions.

Ionic solids generally have high melting and boiling points due to the strong electrostatic attractions between the (+) and (-) ions.

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By contrast, molecular compounds do not conduct electricity either in the solid or the liquid phase. We model these compounds as substances in which the primary building block is a neutral molecule. In each molecule, the electrons are held tightly by the cores and are unable to move and conduct electricity (raisins do not change bowls). When molecular solids are melted, they break into their constituent molecules, which are electrically neutral, and their motion cannot conduct electricity. Many molecular solids also form non-conducting solutions when they dissolve in water (exception – acids and bases which react with water to form ions).
The forces between the electrons and cores within molecules are like those between the T and B tapes – some electrons are simultaneously attracted to more than one core, resulting in a strong electrical interactions between the atoms (bonds). The attractions between the molecules are somewhat like those between either top or bottom tape and the strip of paper – slightly uneven distributions of the mobile negative charge within molecules give rise to relatively weak attractions between the molecules.

Hence, molecular compounds have relatively low melting and boiling points. As a general rule, the smaller the molecule, the weaker these attractive forces (called London forces) and the lower the melting and boiling points. As an example you can cite the halogens. The lighter molecules of fluorine and chlorine are gases at room temperature (i.e., room temperature is above their bp’s); bromine is a liquid (room temperature is greater than its mp, but lower than its bp) and iodine is a solid (room temperature is lower than its mp). This point may seem trivial but it is worth making because unless it is called to their attention, most students do not make the connection between mp and bp and the phase at a given temperature. Instead, students think substances are solid, liquid or gas because that is their nature.

9. Ionic nomenclature – worksheet 3

This is the appropriate time to introduce conventions for distinguishing ionic compound from molecular compounds by nomenclature. Because ions combine to form electrically neutral compounds, the names for ionic compounds do not contain cues as to the number of each kind of ion present. This system presumes that one knows the formula and charge of the constituent ions. This poses a problem for students until they memorize a few rules about the charges of ions, making use of the Periodic Table.

1. Metals form (+) ions and monatomic with the exception of mercury(I) – Hg$^{2+}$. The charges on ions formed from elements in groups 1, 2 and the upper part of 3 [13] is the same as the group number; e.g., the alkali metals form +1 ions, the alkaline earths form +2 ions, aluminum is +3. The ion name is the same as that of the element.

2. With the exception of Cd, Zn and Ag, which have only one charge, the rest of the metals can form ions with multiple charges. These are distinguished by a Roman numeral after the ion name; e.g., Fe$^{2+}$ is iron(II) and Fe$^{3+}$ is iron(III), but zinc is always Zn$^{2+}$.

3. Non-metals form (-) ions, with the exception of hydrogen, which forms both positive and negative ions. The charge on these ions can be predicted by their position on the periodic table. The halogens are one place away from the noble gases, so they form 1- ions, the oxygen family (chalcogens) are two spots away from the noble gases, so they form 2- ions. These monatomic ions have the –ide ending; e.g., O$^{2-}$ is oxide, P$^{3-}$ is phosphide, etc.

4. Certain elements form polyatomic ions that contain oxygen atoms. There are a few general rules that can help students learn these (ions with the –ate ending contain one more oxygen atom than ones with the –ite ending), but the patterns are not obvious.

If you can project documents from your computer, you might find the website Ionic Nomenclature Help11 helpful in this discussion. Until students master the names, formulas and charges of the most common of these ions, they will have difficulty writing names from formulas and formulas from names. They find it difficult to distinguish between sulfur trioxide and the sulfite ion. You will find that unless they are required to demonstrate mastery of some of the most common ions, they will refer to ionic compounds by their formulas, “Please pass the em-gee-see-el-two.” rather than by their names (magnesium chloride). Worksheet 3 gives them the opportunity to practice this skill.

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11 [http://www.mpcfaculty.net/mark_bishop/ionic_nomenclature_help.htm](http://www.mpcfaculty.net/mark_bishop/ionic_nomenclature_help.htm)
10. **Molecular nomenclature – worksheet 4**  
This is a reasonable time to relate the formulas to the names of molecular solids. We employ Greek prefixes to indicate the number of each kind of atom in binary molecular compounds (nomenclature of more complex molecular compounds will not be addressed at this time) Worksheet 4 gives students the opportunity to practice nomenclature.

11. **Quiz**  
This quiz tests students’ mastery of the skills they practiced in worksheet 3.

12. **Representing ions and empirical formulas – worksheet 5**  
Students have difficulty interpreting the meaning of the subscripts in chemical formulas. Their naïve conceptions can lie hidden unless you probe them by requiring students to draw diagrams of ions and the simplest unit of the compound represented by an empirical formula. For example, it is not uncommon to see a student sketch 4 groups of a sulfur atom and an oxygen (\[(SO)_{4}\]) to represent \(SO_4\). Worksheet 5 requires students to write ionic compounds from formulas or break ionic compounds up into their constituent ions and sketch particle diagrams consistent with their formulas. It is important to point out that ionic compounds form crystal lattices, not individual molecules. So the diagram represents a simple cluster of ions that represents a formula unit.

13. **Worksheet 6 (optional) – more practice with empirical formulas**

14. **Unit 6 Review**  
To save time, the first two parts of the nail lab that serves to introduce Unit 7 should be done at the same time as the Unit 6 review. See the Teacher Notes document for Unit 7 for details.

15. **Unit 6 Test**