Properties of Solids
Structure and Bonding

Introduction

Looking for patterns in the properties of different substances can help us understand how and why atoms join together to form compounds. What kinds of forces hold atoms together? How does the nature of the forces holding atoms together influence the properties of a material?

Concepts

- Chemical bonds
- Ionic bonding
- Covalent bonding
- Metallic bonding

Background

Groups of atoms are held together by attractive forces that we call *chemical bonds*. The origin of chemical bonds is reflected in the relationship between force and energy in the physical world. Think about the force of gravity—in order to overcome the force of attraction between an object and the Earth, we have to supply energy. Whether we climb a mountain or throw a ball high into the air, we have to supply energy. Similarly, in order to break a bond between two atoms, energy must be added to the system, usually in the form of heat, light or electricity. The opposite is also true: whenever a bond is formed, energy is released.

The term *ionic bonding* is used to describe the attractive forces between oppositely charged ions in an ionic compound. An ionic compound is formed when a metal reacts with a non-metal to form positively charged cations and negatively charged anions, respectively. The oppositely charged ions arrange themselves in a tightly packed, extended three-dimensional structure called a crystal lattice (see Figure 1). The net attractive forces between oppositely charged ions in the crystal structure are called ionic bonds.

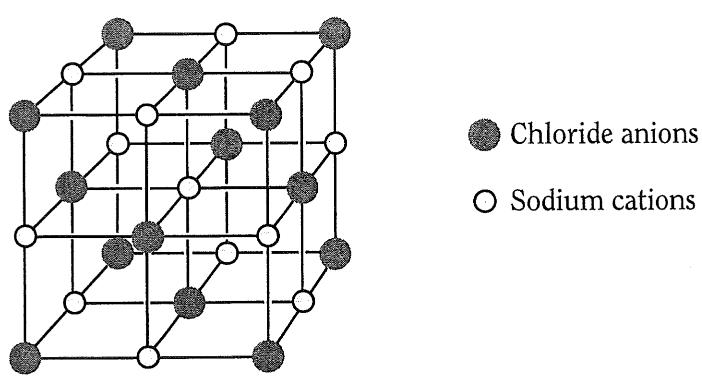

Figure 1. Crystal Structure of Sodium Chloride.

Covalent bonding represents another type of attractive force between atoms. Covalent bonds are defined as the net attractive forces resulting from pairs of electrons that are shared between atoms (the shared electrons are attracted to the nuclei of both atoms in the bond). A group of atoms held together by covalent bonds is called a molecule. Atoms may share one, two or three pairs of electrons between them to form single, double, and triple bonds, respectively.

Substances held together by covalent bonds are usually divided into two groups based on whether individual (distinct) molecules exist or not. In a *molecular solid,* individual molecules in the solid state are attracted to each other by relatively weak intermolecular forces between the molecules. *Covalent-network solids,* on the other hand, consist of atoms forming covalent bonds with each other in all directions. The result is an almost infinite network of strong covalent bonds—there are no individual molecules.

Covalent bonds may be classified as polar or nonpolar. The element chlorine, for example, exists as a diatomic molecule, Cl_2. The two chlorine atoms are held together by a single covalent bond, with the two electrons in the bond equally shared between the two identical chlorine atoms. This type of bond is called a *nonpolar* covalent bond. The compound hydrogen chloride (HCl) consists of a hydrogen atom and a chlorine atom that also share a pair of electrons between them. Because the two atoms are different, however, the electrons in the bond are not equally shared between the atoms. Chlorine has a greater *electronegativity* than hydrogen—it attracts the bonding electrons more strongly than hydrogen. The covalent bond between hydrogen and chlorine is an example of a *polar* bond. The distribution of bonding electrons in a nonpolar versus polar bond is shown in Figure 2. Notice that the chlorine atom in HCl has a partial negative charge (δ^-) while the hydrogen atom has a partial positive charge (δ^+).

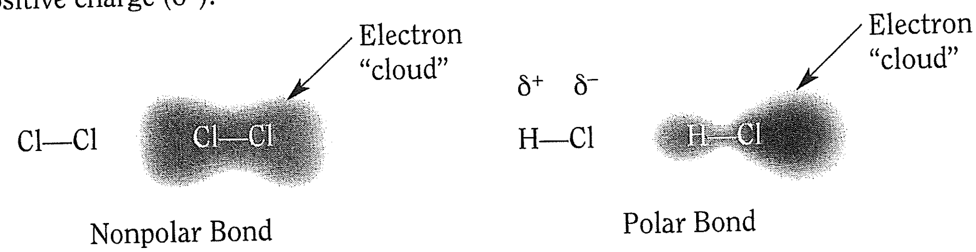

Figure 2. Nonpolar versus Polar Covalent Bonds.

The special properties of metals compared to nonmetals reflect their unique structure and bonding. Metals typically have a small number of valence electrons available for bonding. The valence electrons appear to be free to move among all of the metal atoms, which must exist therefore as positively charged cations. *Metallic bonding* describes the attractive forces that exist between closely packed metal cations and free-floating valence electrons in an extended three-dimensional structure.

Experiment Overview

The purpose of this experiment is to study the physical properties of common solids and to investigate the relationship between the type of bonding in a substance and its properties. The following physical properties will be studied:

- Volatility and Odor: Volatile substances evaporate easily and may have an odor.
- Melting Point: The temperature at which a solid turns into a liquid.
- Solubility: Ability of one substance to dissolve in another. Water is a highly polar solvent. Hexane is nonpolar.
- Conductivity: Ability to conduct electricity.
- Hardness: Resistance of a substance to being scratched.
- Brittleness: Tendency of a solid to break or crumble when a stress is applied.

Pre-Lab Questions

1. A student wanted to illustrate the structure of magnesium chloride and decided simply to replace the Na^+ ions in Figure 1 with Mg^{2+} ions. What would be wrong with the resulting picture?

2. Covalent bonds may be classified as polar or nonpolar based on the difference in electronegativity between two atoms. Look up electronegativity values in your textbook:

   (a) Why are C—H bonds considered nonpolar?

   (b) Which is more polar, an O—H or N—H bond?

3. The three dimensional structure of diamond, a crystalline form of the element carbon, is shown in Figure 3. Use this structure to explain why diamond is the hardest known material.

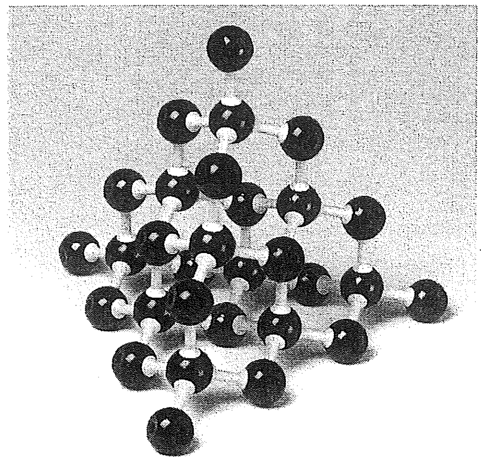

Figure 3.

Materials

Aluminum shot or granules, Al, 0.5 g
Distilled water and wash bottle
Hexane, C_6H_{14}, 5 mL
Silicon dioxide (sand), SiO_2, 0.2–0.3 g
Sodium chloride (salt), NaCl, 0.2–0.3 g
Stearic acid, $C_{18}H_{36}O_2$, 0.2–0.3 g
Sucrose (sugar), $C_{12}H_{22}O_{11}$, 0.2–0.3 g
Minerals for hardness testing (optional)
  Aluminum strip
  Candle (paraffin wax)
  Halite or rock salt (sodium chloride)
  Quartz (silicon dioxide)
  Rock candy (sucrose)
Penny and nail to test hardness (optional)
Balance, centigram
Weighing dishes, 5

Aluminum evaporating dish or Pyrex® watch glass
Beaker, 150-mL
Boiling stones
Bunsen burner*
Conductivity tester, low-voltage*
Hot plate*
Mortars and pestles, 5*
Pipets, Beral-type, 2
Reaction plate, 24-well
Spatula
Stirring rod or toothpicks
Test tubes, Pyrex®, small, 5, or Ceramic spot plate
Test tube rack
Test tube holder (clamp)

*Students may share conductivity testers, hot plates, laboratory burners, and mortars and pestles.

Safety Precautions

Hexane is a flammable organic solvent and a dangerous fire risk. Keep away from flames, heat, and other sources of ignition. Cap the solvent bottle and work with hexane in a fume hood or designated work area well away from the Bunsen burner used in step 12. Avoid contact of all chemicals with eyes and skin. Wear chemical splash goggles and chemical-resistant gloves and apron. Wash hands thoroughly with soap and water before leaving the lab.

ocedure

1. Prepare a boiling water bath for use in step 11: Half-fill a 150-mL beaker with water, add a boiling stone, and heat the beaker on a hot plate at a medium setting.

2. Label five weighing dishes and obtain 0.2–0.3 g samples of each solid in the appropriate weighing dish. Record the color and appearance of each solid in the data table.

3. Test the volatility and odor of each solid by wafting any vapors to your nose with your hand. Record all observations in the data table.

4. Test the conductivity of each solid by touching the wires of the conductivity tester directly to the solid. Record the conductivity of each sample in the data table.

5. Obtain a 24-well reaction plate and add a *small* amount of each solid (about the size of a grain of rice) to separate wells A1–A5, in the order shown in the data table.

6. Add about 20 drops of water to each well. Stir each mixture and observe whether the solid dissolves in water. Record the solubility (soluble, partially soluble, or insoluble) in the data table.

7. *For water-soluble substances only:* Determine the conductivity of the aqueous solution by placing the wires directly into the liquid. Record the results in the data table.

8. Label five small test tubes or a ceramic spot plate and add a small amount of each solid, about the size of a grain of rice, to separate test tubes.

9. Add about 20 drops of hexane to each test tube. Stir each mixture and observe whether the solid dissolves in hexane. Record the results in the data table.

10. Obtain a large, disposable aluminum evaporating dish or Pyrex® watch glass and place a small, pea-sized amount of each solid in separate locations on the dish.

11. Set the dish on top of the boiling water bath and heat the solids for 1–2 minutes. Observe whether any of the solids melt and record the observations in the data table.

12. *For solids that did not melt at the boiling water bath temperature:* Place a small, pea-sized amount of each solid in a clean and dry, Pyrex® test tube. Using a test tube holder, heat the test tube in a burner flame for 1–2 minutes. Record observations in the data table.

13. Test the brittleness of each solid by placing a small sample in the mortar designated for it and grinding with the pestle. Record the observations in the data table.

14. *(Optional)* Test the hardness of the mineral samples by trying to scratch them with a fingernail, a penny, and a nail. Record observations on the data sheet.

Name: ______________________

Class/Lab Period: ______________________

Properties of Solids

Data Table

Physical Property	Aluminum	Silicon Dioxide	Sodium Chloride	Stearic Acid	Sucrose
Color and Appearance					
Volatility and Odor					
Conductivity (Solid)					
Solubility in Water					
Conductivity of Aqueous Solution					
Solubility in Hexane					
Brittleness					
Melting Point*					

*The average temperature of a Bunsen burner flame is greater than 1000 °C.

(Optional) Use this space to write down your observations of the hardness of mineral samples.

-L Questions *(Use a separate sheet of paper to answer the following questions.)*

1. Compare the volatility and odor of stearic acid and sucrose. Which is more volatile? Why? Is it possible for a compound to be volatile but have no odor? Explain.
2. Both stearic acid and sucrose are molecular substances, but one is polar and the other is nonpolar. Compare the solubility of the two compounds in water and in hexane to determine which is which.
3. Based on the answers to Questions #1 and 2, predict whether the intermolecular forces (forces between molecules) are stronger in polar or nonpolar substances.
4. In order for a substance to conduct electricity, it must have free-moving charged particles.
   (a) Explain the conductivity results observed for sodium chloride in the solid state and in aqueous solution.
   (b) Would you expect molten sodium chloride to conduct electricity? Why or why not?
   (c) Use the model of metallic bonding described in the *Background* section to explain why metals conduct electricity.
5. Name the three hardest substances that were tested. To what classes of solids do these substances belong? What general feature do these three types of solids have in common?
6. Compare the hardness and brittleness of aluminum versus salt. Suggest a reason, based on the crystal structure of metals versus ionic compounds, why hardness and brittleness are not the same thing.
7. Complete the following table (some of the entries have been filled in for you):

General Properties	**Type of Solid**			
	Covalent-network	**Ionic**	**Metallic**	**Molecular**
Melting Point			Low to high	
Solubility				Depends on polarity
Conductivity of Solid		Nonconductors		
Hardness	Very hard			
Brittleness				