Qualitative Analysis and Chemical Bonding

Introduction

Looking for patterns in the properties of solids can help us understand how and why atoms join together to form compounds. What kinds of forces hold atoms together? How do these forces influence the properties of materials? Use your knowledge of the relationship between chemical bonding type and the properties of substances to determine the identity of mystery solids.

Concepts

- Chemical bonds
- Covalent bonding
- Ionic bonding
- Metallic bonding
- Physical and chemical properties
- Electronegativity

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. 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* describes attractive forces between oppositely charged ions in an ionic compound. An ionic compound is formed when a metal reacts with a nonmetal to form positively charged cations and negatively charged anions, respectively. The oppositely charged ions are arranged 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.

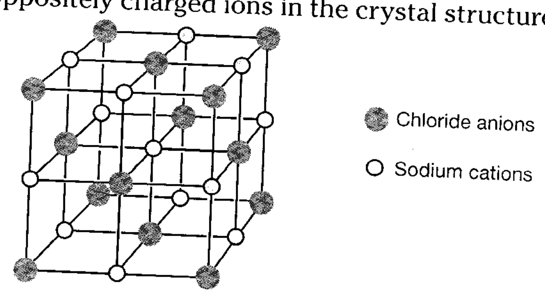

Figure 1. Crystal Structure of Sodium Chloride.

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.

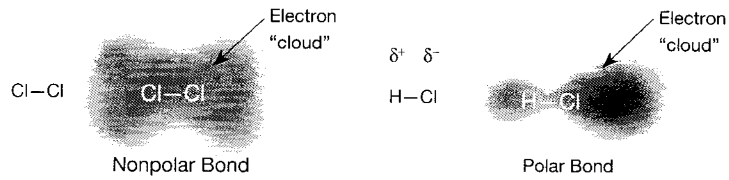

Figure 2. Nonpolar versus Polar Covalent Bonds.

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 shared between 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^+).

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, some of which must exist or act 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 advanced inquiry lab is to identify twelve unknown solids based on systematic testing of their physical and chemical properties. The lab begins with an introductory activity to select measurable properties that will help identify the type of bonding in a solid. Given four solids representing the four types of chemical bonds—ionic, polar covalent, nonpolar covalent and metallic—students review the properties of each solid with a minimum of four tests. The results provide a basis for a guided-inquiry design of a flow chart procedure to distinguish and identify twelve unknown solids.

Pre-Lab Questions

Solid	Bond Type	Melting Point, °C	Color	Solubility in Water, 25 °C
Sodium iodide, NaI	Ionic	661	White	Soluble
Lactose, $C_{12}H_{22}O_{11}$	Polar covalent	203	White	Soluble
Iodine, I_2	Nonpolar covalent	114	Dark gray	Slightly soluble
Tin, Sn	Metallic	232	Shiny gray	Insoluble

1. Considering the data in the above table, explain the following observations based on the type of chemical bonding and intermolecular forces between atoms, molecules or ions in the solid state.

   a. Iodine has a much lower melting point than sodium iodide.

   b. Iodine is slightly soluble in water, while lactose is very soluble.

2. Predict the properties of the following solids.

Solid	Bond Type	Melting Point High or Low	Solubility in Water, 25 °C
Iodine chloride, ICl			
Cobalt nitrate, $Co(NO_3)_2$			
Sulfur, S_8			
Lead, Pb			

Materials (for each lab group)

Ethyl alcohol, CH_3CH_2OH, 20 mL

Hexane, C_6H_{14}, 20 mL

Hydrochloric acid solution, HCl, 0.1 M, 20 mL

Sodium hydroxide solution, NaOH, 0.1 M, 20 mL

Water, distilled or deionized, 20 mL

Known samples for the *Introductory Activity*

- Copper(II) sulfate, $CuSO_4$, 2 g
- Dextrose, monohydrate, $C_6H_{12}O_6 \cdot H_2O$, 2 g
- Paraffin wax, C_nH_{2n+2} (n = 20–40), 2 g
- Zinc, Zn, 2 g

Unlabeled samples, 6, 3 g each

Aluminum dish

Beaker, 100-mL

Bunsen burner

Conductivity meter or tester

Hot plate

pH paper

Stirring rod

Test tube rack

Test tube holder

Test tubes, 6

Thermometer

Tongs

Safety Precautions

Hexane and ethyl alcohol are flammable organic solvents and dangerous fire risks. Keep away from flames, heat, and other sources of ignition. Cap the solvent bottles and work with hexane and ethyl alcohol in a fume hood or designated work area. Addition of a denaturant makes ethyl alcohol poisonous; it cannot be made nonpoisonous. Copper(II) sulfate is a skin and respiratory tract irritant and is toxic by ingestion. Graphite powder is a fire and inhalation risk. Dilute sodium hydroxide and acid solutions are irritating to skin and eyes. Avoid contact of all chemicals with eyes and skin. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the lab. Please follow all laboratory safety guidelines.

Introductory Activity

Identifying Properties of Chemical Bonds

1. Four representative chemicals are provided for preliminary testing to identify physical and chemical properties that can be used for development of a qualitative analysis scheme. The chemicals and the type of bonding in each are:

   Copper(II) sulfate—ionic bonding

   Paraffin wax—nonpolar covalent bonding

   Dextrose—polar covalent bonding

   Zinc—metallic bonding

2. Observe the color and appearance of each solid and perform the following qualitative tests on each: (a) solubility in water, solubility in hexane, and solubility in alcohol; (b) high or low melting point; (c) conductivity of solid and conductivity of aqueous solution; (d) pH of solution; (e) reaction with acid (0.1 M HCl) and reaction with base (0.1 M NaOH).

3. Consult the *Materials* list when developing small-scale or microscale procedures for these tests. Note that a total of 2 g of each solid and 20 mL of each solvent are provided for testing.

4. To test the melting point of a substance, first place a small amount of each solid in separate locations in an aluminum evaporating dish. Hold the dish above a boiling water bath with tongs and observe if the solids melt at < 100 °C.
5. If a substance did *not* melt at < 100 °C, place a pea-size amount of solid in a borosilicate glass test tube. Heat the test tube in a medium burner flame for 1–2 minutes. Record observations. **Note:** To test the melting point in a medium burner flame, start with a fuel-rich yellow flame and adjust the air inlet until all yellow just disappears. The result is a light blue flame without an inner cone; the temperature is < 500 °C.
6. Record the results of qualitative testing in a data table.

Guided-Inquiry Design and Procedure

Development of a Qualitative Analysis Scheme

A school in the district has discovered a set of chemicals that are missing labels. The science teacher has recovered the potential missing labels, but needs to match them with the correct bottles. The AP Chemistry class has been asked to design a procedure to identify the 12 unknown chemicals. You may use any series of tests that deal with the properties of solids.

Form a working group with other students and discuss the following questions.

1. Compare your group's results from the *Introductory Activity* with those of other groups. From these discussions and your test data, list general physical properties that can be associated with each type of bonding in a solid.
2. Using yes–no logic, create a flow chart that can be used to characterize an unknown solid as ionic, polar covalent, nonpolar covalent or metallic.
3. If given a white solid, what testing results would help you to identify the solid as polar covalent?
4. Each group will be given six unknown solids to evaluate. Using the flow chart as a guide, write a detailed, step-by-step procedure for testing the solids and identifying the type of bonding in each. Include the materials and equipment that will be needed, safety precautions that must be followed, amounts of chemicals to use, etc.
5. Carry out the flow chart tests on the six unknowns and record the results of each test in an appropriate data table.
6. Identify each solid as ionic, polar covalent, nonpolar covalent or metallic.
7. Share your group's data with a group that analyzed the other six unknowns.
8. Copies of the missing labels have been assembled on a sheet of paper. Compare the results of testing the unknowns with the chemical labels. As a class, match the labels with as many unknowns as possible.
9. Perform any additional tests as needed to verify the identity of any remaining chemicals whose identities are ambiguous.

Opportunities for Inquiry

The chemistry teacher suspects vandalism has occurred in the stockroom of the high school. An order of chemicals arrived and was placed in the storeroom over the weekend. When the chemistry teacher returned on Monday, the labels on six identically sized bottles appeared misaligned, as if they had been switched. You are asked to use your flow chart to verify or refute the identity of the solids in the six bottles.

AP Chemistry Review Questions

Integrating Content, Inquiry and Reasoning

1. 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? Explain.

2. To convert the following compounds from a solid to a liquid, what types of intermolecular forces must be overcome?
   a. $I_2(s) \rightarrow I_2(l)$

   b. $H_2O(s) \rightarrow H_2O(l)$

   c. $NaI(s) \rightarrow NaI(l)$

   d. $C_{16}H_{32}(s) \rightarrow C_{16}H_{32}(l)$

3. In order for a substance to conduct electricity, it must have free-moving charged particles.
   a. Explain the conductivity results observed for ionic compounds 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.