

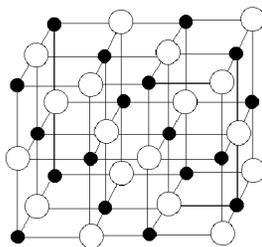
Guard well your spare moments...
They are like uncut diamonds.
Discard them and their value will never
be known. Improve them and they will
become the brightest gems in a useful
life. ---Ralph Waldo Emerson

Classifying Solids

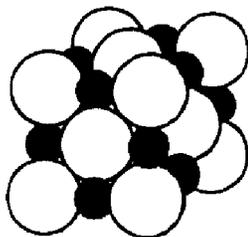
The behavior of substances in the solid state is governed mainly by the way in which the atoms, ions or molecules of the substance are arranged and the forces that hold them together. Although there are many different substances, the types of solids can be roughly divided into 4 categories:

- molecular
- ionic
- network covalent (or macromolecular)
- metallic

Each of these is distinguished by the kinds of particles that make up the three-dimensional structure of the solid. If the structure has a regular pattern then the solid is described as *crystalline*. A familiar example would be sodium chloride, NaCl, in which sodium and chlorine ions alternate in an extended structure known as a *space lattice*:



In this structure each ion occupies an imaginary point in space called a *lattice point*. This is an example of an ionic solid since ions are the structural unit. A more realistic representation of the arrangement of ions in the crystal would look more like the structure below:



Any crystalline solid would exhibit a similar orderly arrangement of particles. Some solids, however, do not have such organization on the molecular level. Substances such as glass, rubber and many plastics are called *amorphous* solids because they lack a typical space-lattice organization. In this experiment we are concerned only with crystalline solids.

Molecular crystals consist of either molecules or atoms in orderly arrangements. These arrangements are maintained by weak attractive forces such as hydrogen bonding, dipole forces and dispersion forces. Because of this, such solids generally have low to medium melting points (below about 400°C).

Since the individual units that make up the molecular crystal are neutral in charge and cannot move appreciably in the solid state, these solids are poor electrical conductors. Even when melted, electrical conductivity is negligible. And if the solids dissolve in water (see below) there is still very little, if any, conductivity due to lack of charges on the molecules or atoms.

Some molecular crystals will dissolve in water (if the units of the crystal are polar) while others which are composed of non-polar molecules or atoms may dissolve in non-polar solvents such as hexane. (recall the solubility of the non-polar halogens in hexane).

Most organic compounds (composed chiefly of C, H and O) form molecular crystals. While these generally melt at low temperatures (less than 250°C) they often decompose or char before melting.

Ionic crystals are, of course, composed of ions. These are held tightly in place in the space-lattice by strong electrostatic forces because of the opposite charges of the ions. These very strong forces contribute to high melting points, although some hydrated ionic compounds may appear to "melt" at lower temperatures (see below).

Although the ions are charged, they cannot move in the solid and so these crystals do not conduct electricity. However, when melted, most of the attractions have been overcome and molten ionic compounds conduct very well. Most ionic compounds also dissolve to at least some extent in water. The amount that dissolves is completely separated into ions and thus makes an electrolyte solution which is a good conductor.

Hydrated ionic compounds contain water molecules at some of the lattice points. These molecules are held in place by the attraction of the polar water molecules for the charged ions. However, heating will overcome these attractions and release the water (the crystal structure generally collapses or at least changes). Sometimes if the water is released slowly by gentle heating it can dissolve the ionic solid and make it appear as if it is melting at a low temperature.

Network covalent crystals (or macromolecular solids) are held together by a network of covalent bonds that extends throughout the solid sample. Diamond is a classic example in which each atom of carbon is attached to four other carbon atoms. This extensive framework of covalent bonds makes these solids very hard. Also, since each bond must be broken to melt the solid, melting points are typically very high (higher than 1000°)

Network crystals do not conduct electricity because they lack movable charges, and their tight covalent structure makes them insoluble in essentially everything.

Metallic crystals consist of metal cations (minus valence electrons) positioned at the lattice points of the crystal. The valence electrons are given up to a common "soup" or "sea" which flows more or less freely throughout the sample. The lack of directionality in metallic bonding makes most metals bendable and rather soft (certainly compared to a diamond!). Metals are excellent conductors because of all the mobile electrons. This conductivity exists in both the solid and liquid state. The melting points of metals vary widely from mercury at -38°C to tungsten at 3410°C. The effects that contribute to the wide-ranging melting points of metals are beyond the scope of this discussion.

Metals are essentially insoluble in common solvents, although they undergo vigorous chemical reactions with a variety of substances in which they become oxidized and move into aqueous solution as ions. And, of course, metals are recognizable by their optical properties which most other solids lack: lustre and shininess.

Preparing to experiment

You will be provided with the following materials:

1. four solid unknowns
2. heavy aluminum foil
3. 3 small test tubes
4. a 24-well microplate
5. conductivity device
6. hexane

Design an experiment to identify each solid unknown as either molecular (polar), molecular (non-polar), ionic, hydrated ionic, or network covalent.

Study the **Technique** section carefully and review the procedures outlined there before laying out your experiments. Organizing tables in your notebook will probably be very helpful in planning how to proceed.

Technique

1. estimating melting point range

The maximum temperature you can get from a bunsen burner is about 400°C. You can determine whether a solid melts above or below this temperature by placing a **small** amount (smaller than a match head) of the solid on a square of aluminum foil placed on the wire screen for your tripod. All of the solids can be tested simultaneously. Heat from below gently at first, then more strongly and observe the behavior of the solids carefully. Discard the foil square when it has cooled.

2. checking for water of hydration (ionic solids only)

Solids suspected of being ionic can be checked for hydration by heating a **small** amount in a test tube and looking for signs of water condensation on the upper part of the test tube. Heat *gently*.

3. solubility in hexane (non-polar solvent)

CAUTION!!!! Hexane is flammable. No flames should be used in the laboratory when hexane is in use. Hexane will also attack the 24-well plates and should only be used in glass test tubes. Use a minimal amount (1-2 mL) and a few crystals of solid. Discard used hexane in one of the containers found in the fume hoods.

Analysis

Prepare a table of the solids you have tested. Refer to each by its unknown number. Indicate the kind of solid you believe it to be and give a short summary of the experimental evidence that led you to your conclusion. Contradictory observations should be explained in more detail following the table.