Metallic Solids

Metals are characterized by their ability to reflect light, called luster, their high electrical and thermal conductivity, their high heat capacity, and their malleability and ductility. Every lattice point in a pure metallic element is occupied by an atom of the same metal. The packing efficiency in metallic crystals tends to be high, so the resulting metallic solids are dense, with each atom having as many as 12 nearest neighbors.

Bonding in metallic solids is quite different from the bonding in the other kinds of solids we have discussed. Because all the atoms are the same, there can be no ionic bonding, yet metals always contain too few electrons or valence orbitals to form covalent bonds with each of their neighbors. Instead, the valence electrons are delocalized throughout the crystal, providing a strong cohesive force that holds the metal atoms together.

![Image of metallic bonding](image)

*Figure\(\PageIndex{1}\): The "sea of electrons" is free to flow about the crystal of positive metal ions. Image used with permission from Wikipedia.*

Valence electrons in a metallic solid are delocalized, providing a strong cohesive force that holds the atoms together.

The strength of metallic bonds varies dramatically. For example, cesium melts at 28.4°C, and mercury is a liquid at room temperature, whereas tungsten melts at 3680°C. Metallic bonds tend to be weakest for elements that have nearly empty (as in Cs) or nearly full (Hg) valence subshells, and strongest for elements with approximately half-filled valence shells (as in W). As a result, the melting points of the metals increase to a maximum around group 6 and then decrease again from left to right across the d block. Other properties related to the strength of metallic bonds, such as enthalpies of fusion, boiling points, and hardness, have similar periodic trends.

A somewhat oversimplified way to describe the bonding in a metallic crystal is to depict the crystal as consisting of positively charged nuclei in an electron sea (Figure \(\PageIndex{2}\)). In this model, the valence electrons are not tightly bound to any one atom but are distributed uniformly throughout the structure. Very little energy is needed to remove electrons from a solid metal because they are not bound to a single nucleus. When an electrical potential is applied, the
electrons can migrate through the solid toward the positive electrode, thus producing high electrical conductivity. The ease with which metals can be deformed under pressure is attributed to the ability of the metal ions to change positions within the electron sea without breaking any specific bonds. The transfer of energy through the solid by successive collisions between the metal ions also explains the high thermal conductivity of metals. This model does not, however, explain many of the other properties of metals, such as their metallic luster and the observed trends in bond strength as reflected in melting points or enthalpies of fusion.

![Electron-Sea Model of Bonding in Metals](image)

**Figure** \( \PageIndex{2} \): The Electron-Sea Model of Bonding in Metals. Fixed, positively charged metal nuclei from group 1 (a) or group 2 (b) are surrounded by a “sea” of mobile valence electrons. Because a group 2 metal has twice the number of valence electrons as a group 1 metal, it should have a higher melting point.

The general order of increasing strength of interactions in a solid is:

molecular solids < ionic solids ≈ metallic solids < covalent solids.

Some general properties of the four major classes of solids are summarized in Table \( \PageIndex{2} \).

<table>
<thead>
<tr>
<th>Ionic Solids</th>
<th>Molecular Solids</th>
<th>Covalent Solids</th>
<th>Metallic Solids</th>
</tr>
</thead>
<tbody>
<tr>
<td>poor conductors of heat and electricity</td>
<td>poor conductors of heat and electricity</td>
<td>poor conductors of heat and electricity*</td>
<td>good conductors of heat and electricity</td>
</tr>
<tr>
<td>relatively high melting point</td>
<td>low melting point</td>
<td>high melting point</td>
<td>melting points depend strongly on electron configuration</td>
</tr>
<tr>
<td>hard but brittle; shatter under stress</td>
<td>soft</td>
<td>very hard and brittle</td>
<td>easily deformed under stress; ductile and malleable</td>
</tr>
<tr>
<td>relatively dense</td>
<td>low density</td>
<td>low density</td>
<td>usually high density</td>
</tr>
</tbody>
</table>

*Many exceptions exist. For example, graphite has a relatively high electrical conductivity within the carbon planes, and diamond has the highest thermal conductivity of any known substance.
Ionic Solids | Molecular Solids | Covalent Solids | Metallic Solids
--- | --- | --- | ---
Dull surface | Dull surface | Dull surface | Lustrous

*Many exceptions exist. For example, graphite has a relatively high electrical conductivity within the carbon planes, and diamond has the highest thermal conductivity of any known substance.

Example \(\PageIndex{1}\)

Classify Ge, RbI, \(\text{C}_6\)(\text{CH}_3)\text{6}\), and Zn as ionic, molecular, covalent, or metallic solids and arrange them in order of increasing melting points.

**Given:** compounds

**Asked for:** classification and order of melting points

**Strategy:**

A. Locate the component element(s) in the periodic table. Based on their positions, predict whether each solid is ionic, molecular, covalent, or metallic.

B. Arrange the solids in order of increasing melting points based on your classification, beginning with molecular solids.

**Solution:**

A. Germanium lies in the p block just under Si, along the diagonal line of semimetallic elements, which suggests that elemental Ge is likely to have the same structure as Si (the diamond structure). Thus Ge is probably a covalent solid. RbI contains a metal from group 1 and a nonmetal from group 17, so it is an ionic solid containing Rb\(^+\) and I\(^-\) ions. The compound \(\text{C}_6\)(\text{CH}_3)\text{6}\) is a hydrocarbon (hexamethylbenzene), which consists of isolated molecules that stack to form a molecular solid with no covalent bonds between them. Zn is a d-block element, so it is a metallic solid.

B. Arranging these substances in order of increasing melting points is straightforward, with one exception. We expect \(\text{C}_6\)(\text{CH}_3)\text{6}\) to have the lowest melting point and Ge to have the highest melting point, with RbI somewhere in between. The melting points of metals, however, are difficult to predict based on the models presented thus far. Because Zn has a filled valence shell, it should not have a particularly high melting point, so a reasonable guess is \(\text{C}_6\)(\text{CH}_3)\text{6} \sim \text{Zn} \sim \text{RbI} \sim \text{Ge}.\) The actual melting points are \(\text{C}_6\)(\text{CH}_3)\text{6}, 166°C; Zn, 419°C; RbI, 642°C; and Ge, 938°C. This agrees with our prediction.

Exercise \(\PageIndex{1}\)

Classify \(\text{C}_60\), BaBr\(_2\), GaAs, and AgZn as ionic, covalent, molecular, or metallic solids and then arrange them in order of increasing melting points.

**Answer:** \(\text{C}_60\) (molecular) \(<\ AgZn\) (metallic) \(<\ \text{BaBr}_2\) (ionic) \(<\ \text{GaAs}\) (covalent). The actual melting points are \(\text{C}_60\), about 300°C; AgZn, about 700°C; BaBr\(_2\), 856°C; and GaAs, 1238°C.

---

**Summary**