Chapter 1

1. Chapter 1: The Chemical World
2. 1.1: The Scope of Chemistry
3. 1.2: Chemicals Compose Ordinary Things
4. 1.3: Hypothesis, Theories, and Laws
5. 1.4: The Scientific Method: How Chemists Think
6. 1.5: A Beginning Chemist: How to Succeed

• Chapter 2

1. Chapter 2: Measurement and Problem Solving
2. 2.1: Taking Measurements
3. 2.2: Scientific Notation: Writing Large and Small Numbers
4. 2.3: Significant Figures: Writing Numbers to Reflect Precision
5. 2.4: Significant Figures in Calculations
6. 2.5: The Basic Units of Measurement
7. 2.6: Problem Solving and Unit Conversions
8. 2.7: Solving Multistep Conversion Problems
9. 2.8: Units Raised to a Power
10. 2.9: Density
11. 2.10: Numerical Problem-Solving Strategies and the Solution Map
12. 2.E: Measurement and Problem Solving (Exercises)

• Chapter 3

1. Chapter 3: Matter and Energy
2. 3.1: In Your Room
3. 3.2: What is Matter?
4. 3.3: Classifying Matter According to Its State: Solid, Liquid, and Gas
5. 3.4: Classifying Matter According to Its Composition
6. 3.5: Differences in Matter: Physical and Chemical Properties
7. 3.6: Changes in Matter: Physical and Chemical Changes
8. 3.7: Conservation of Mass: There is No New Matter
9. 3.8: Energy
10. 3.9: Energy and Chemical and Physical Change
11. 3.10: Temperature: Random Motion of Molecules and Atoms
12. 3.11: Temperature Changes: Heat Capacity
13. 3.12: Energy and Heat Capacity Calculations
14. 3.E: Exercises
• Chapter 4
  1. **Chapter 4: Atoms and Elements**
  2. 4.1: Experiencing Atoms at Tiburon
  3. 4.2: Indivisible: The Atomic Theory
  4. 4.3: The Nuclear Atom
  5. 4.4: The Properties of Protons, Neutrons, and Electrons
  6. 4.5: Elements: Defined by Their Numbers of Protons
  7. 4.6: Looking for Patterns: The Periodic Law and the Periodic Table
  8. 4.7: Ions: Losing and Gaining Electrons
  9. 4.8: Isotopes: When the Number of Neutrons Varies
  10. 4.9: Atomic Mass: The Average Mass of an Element’s Atoms

• Chapter 5
  1. **Chapter 5: Molecules and Compounds**
  2. 5.1: Sugar and Salt
  3. 5.2: Compounds Display Constant Composition
  4. 5.3: Chemical Formulas: How to Represent Compounds
  5. 5.4: A Molecular View of Elements and Compounds
  6. 5.5: Writing Formulas for Ionic Compounds
  7. 5.6: Nomenclature: Naming Compounds
  8. 5.7: Naming Ionic Compounds
  9. 5.8: Naming Molecular Compounds
  10. 5.9: Naming Acids
  11. 5.10: Nomenclature Summary
  12. 5.11: Formula Mass: The Mass of a Molecule or Formula Unit

• Chapter 6
  1. **Chapter 6: Chemical Composition**
  2. 6.1: How Much Sodium?
  3. 6.2: Counting Nails by the Pound
  4. 6.3: Counting Atoms by the Gram
  5. 6.4: Counting Molecules by the Gram
  6. 6.5: Chemical Formulas as Conversion Factors
  7. 6.6: Mass Percent Composition of Compounds
  8. 6.7: Mass Percent Composition from a Chemical Formula
  9. 6.8: Calculating Empirical Formulas for Compounds
  10. 6.9: Calculating Molecular Formulas for Compounds

• Chapter 7
4. 10.3: Lewis Structures of Ionic Compounds: Electrons Transferred
5. 10.4: Covalent Lewis Structures: Electrons Shared
6. 10.5: Writing Lewis Structures for Covalent Compounds
7. 10.6: Resonance: Equivalent Lewis Structures for the Same Molecule
8. 10.7: Predicting the Shapes of Molecules
9. 10.8: Electronegativity and Polarity: Why Oil and Water Don’t Mix

• Chapter 11
  1. Chapter 11: Gases
  2. 11.1: Extra-Long Straws
  3. 11.2: Kinetic Molecular Theory: A Model for Gases
  4. 11.3: Pressure: The Result of Constant Molecular Collisions
     5. 11.4: Boyle’s Law: Pressure and Volume
     6. 11.5: Charles’s Law: Volume and Temperature
     7. 11.6: Gay-Lussac's Law: Temperature and Pressure
  8. 11.7: The Combined Gas Law: Pressure, Volume, and Temperature
  9. 11.8: Avogadro’s Law: Volume and Moles
10. 11.9: The Ideal Gas Law: Pressure, Volume, Temperature, and Moles
11. 11.10: Mixtures of Gases: Why Deep-Sea Divers Breathe a Mixture of Helium and Oxygen
12. 11.11: Gases in Chemical Reactions

• Chapter 12
  1. Chapter 12: Liquids, Solids, and Intermolecular Forces
     2. 12.1: Interactions between Molecules
     3. 12.2: Properties of Liquids and Solids
     4. 12.3: Intermolecular Forces in Action: Surface Tension and Viscosity
     5. 12.4: Evaporation and Condensation
     6. 12.5: Melting, Freezing, and Sublimation
  7. 12.6: Types of Intermolecular Forces: Dispersion, Dipole–Dipole, Hydrogen Bonding, and Ion-Dipole
  8. 12.7: Types of Crystalline Solids: Molecular, Ionic, and Atomic
     9. 12.8: Water: A Remarkable Molecule

• Chapter 13
  1. Chapter 13: Solutions
     2. 13.1: Prelude - Tragedy in Cameroon
     3. 13.2: Solutions: Homogeneous Mixtures
     4. 13.3: Solutions of Solids Dissolved in Water: How to Make Rock Candy
     5. 13.4: Solutions of Gases in Water: How Soda Pop Gets Its Fizz
     6. 13.5: Solution Concentration: Mass Percent
Learning Objectives

- Explain the following laws within the Ideal Gas Law

We often take a lot of things for granted. We just assume that we will get electric power when we connect a plug to an electrical outlet. The wire that comprises that outlet is almost always copper, a material that conducts electricity well. The unique properties of the solid copper allow electrons to flow freely through the wire and into whatever device we connect it to. Then we can enjoy music, television, work on the computer, or whatever other activity we want to undertake.

Classes of Crystalline Solids

Crystalline substances can be described by the types of particles in them and the types of chemical bonding that takes place between the particles. There are four types of crystals: (1) ionic, (2) metallic, (3) covalent network, and (4) molecular. Properties and several examples of each type are listed in the following table and are described in the table below.

<table>
<thead>
<tr>
<th>Type of Crystalline Solid</th>
<th>Examples (formulas)</th>
<th>Melting Point (°C)</th>
<th>Normal Boiling Point (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ionic</td>
<td>( \text{NaCl} )</td>
<td>801</td>
<td>1413</td>
</tr>
</tbody>
</table>

Table \(\PageIndex{1}\): Crystalline Solids - Melting and Boiling Points
<table>
<thead>
<tr>
<th>Type of Crystalline Solid</th>
<th>Examples (formulas)</th>
<th>Melting Point (°C)</th>
<th>Normal Boiling Point (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Metalic</td>
<td>(\text{CaF}_2)</td>
<td>1418</td>
<td>1533</td>
</tr>
<tr>
<td></td>
<td>(\text{Hg})</td>
<td>-39</td>
<td>630</td>
</tr>
<tr>
<td></td>
<td>(\text{Na})</td>
<td>371</td>
<td>883</td>
</tr>
<tr>
<td></td>
<td>(\text{Au})</td>
<td>1064</td>
<td>2856</td>
</tr>
<tr>
<td></td>
<td>(\text{W})</td>
<td>3410</td>
<td>5660</td>
</tr>
<tr>
<td>Covalent Network</td>
<td>(\text{B})</td>
<td>2076</td>
<td>3927</td>
</tr>
<tr>
<td></td>
<td>(\text{C}) (diamond)</td>
<td>3500</td>
<td>3930</td>
</tr>
<tr>
<td></td>
<td>(\text{Si}_2\text{O}_3)</td>
<td>1600</td>
<td>2230</td>
</tr>
<tr>
<td></td>
<td>(\text{H}_2)</td>
<td>-259</td>
<td>-253</td>
</tr>
<tr>
<td>Molecular</td>
<td>(\text{NH}_3)</td>
<td>114</td>
<td>184</td>
</tr>
<tr>
<td></td>
<td>(\text{H}_2\text{O})</td>
<td>0</td>
<td>100</td>
</tr>
</tbody>
</table>

**Ionic crystals** — The ionic crystal structure consists of alternating positively-charged cations and negatively-charged anions (see figure below). The ions may either be monatomic or polyatomic. Generally, ionic crystals form from a combination of Group 1 or 2 metals and Group 16 or 17 nonmetals or nonmetallic polyatomic ions. Ionic crystals are hard and brittle and have high melting points. Ionic compounds do not conduct electricity as solids, but do conduct when molten or in aqueous solution.

**Metallic crystal** — Metallic crystals consist of metal cations surrounded by a "sea" of mobile valence electrons (see figure below). These electrons, also referred to as delocalized electrons, do not belong to any one atom, but are capable of moving through the entire crystal. As a result, metals are good conductors of electricity. As seen in the table above, the melting points of metallic crystals span a wide range.
Covalent network crystals -- A covalent network crystal consists of atoms at the lattice points of the crystal, with each atom being covalently bonded to its nearest neighbor atoms (see figure below). The covalently bonded network is three-dimensional and contains a very large number of atoms. Network solids include diamond, quartz, many metalloids, and oxides of transition metals and metalloids. Network solids are hard and brittle, with extremely high melting and boiling points. Being composed of atoms rather than ions, they do not conduct electricity in any state.

Molecular crystals -- Molecular crystals typically consist of molecules at the lattice points of the crystal, held together by relatively weak intermolecular forces (see figure below). The intermolecular forces may be dispersion forces in the case of nonpolar crystals, or dipole-dipole forces in the case of polar crystals. Some molecular crystals, such as ice, have molecules held together by hydrogen bonds. When one of the noble gases is cooled and solidified, the lattice points are individual atoms rather than molecules. In all cases, the intermolecular forces holding the particles together are far weaker than either ionic or covalent bonds. As a result, the melting and boiling points of molecular crystals are much lower. Lacking ions or free electrons, molecular crystals are poor electrical conductors.
Some general properties of the four major classes of solids are summarized in Table 4.

<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>
<tr>
<td>dull surface</td>
<td>dull surface</td>
<td>dull surface</td>
<td>lustrous</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.

Example

Classify \( \text{Ge}, \text{RbI}, \text{C}_{6}\text{(CH}_{3}\text{)}_{6}, \text{Zn} \) 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 \(\text{C}_6(\text{CH}_3)_6\) 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 \(\text{C}_6(\text{CH}_3)_6\) 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

\[
\text{C}_6(\text{CH}_3)_6 < \text{Zn} \sim \text{RbI} < \text{Ge}.
\]

The actual melting points are \(\text{C}_6(\text{CH}_3)_6\), 166°C; Zn, 419°C; RbI, 642°C; and Ge, 938°C. This agrees with our prediction. Exercise \(\PageIndex{1}\)

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

Answer

CO\(_2\) (molecular) < AgZn (metallic) ~ BaBr\(_2\) (ionic) < GaAs (covalent).

The actual melting points are CO\(_2\), about -15.6°C; AgZn, about 700°C; BaBr\(_2\), 856°C; and GaAs, 1238°C.

---

Summary

Ionic crystals are composed of alternating positive and negative ions. Metallic crystals consist of metal cations surrounded by a "sea" of mobile valence electrons. Covalent crystals are composed of atoms which are covalently bonded to one another. Molecular crystals are held together by weak intermolecular forces.

Contributions & Attributions

This page was constructed from content via the following contributor(s) and edited (topically or extensively) by the LibreTexts development team to meet platform style, presentation, and quality:

- CK-12 Foundation by Sharon Bewick, Richard Parsons, Therese Forsythe, Shonna Robinson, and Jean Dupon.
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)