Chapter Objectives:
• Learn the differences between the solid, liquid, and gas state, and how the polarity of molecules influences those states.
• Learn the different types of intermolecular forces between molecules.
• Learn about the energy changes that accompany phase changes, and how to read phase diagrams.

Gases vs. Liquids and Solids
• In gases, the particles in the sample are widely separated, because the attractive forces between the particles are very weak.
• In liquids, there are strong intermolecular forces between the particles, which hold them in close contact, while still letting them slip and slide over one another.
• In solids, the intermolecular forces are so strong that the particles are held rigidly in place.
Solids, Liquids, and Gases

The Physical States of Matter

- **Solids**
  - have strong intermolecular forces.
  - have high densities in comparison to gases.
  - are *rigid* (have a definite shape) and *incompressible* (have a definite volume).
  - may be *crystalline* (ordered) [e.g. table salt] or *amorphous* (disordered) [e.g., plastics].
The Physical States of Matter

- **Liquids**
  - have moderate intermolecular forces.
  - have high densities in comparison to gases.
  - are *fluid* (they flow, and have an indefinite shape) and *incompressible* (have a definite volume); they conform to the shape of their containers (they form *surfaces*).

- **Gases**
  - have weak intermolecular forces.
  - have low densities
  - are *fluid* and *compressible*; they have no definite shape or volume, and conform to the container shape, but fill the entire volume (i.e., they do not form surfaces).
Chapter 10 Liquids and Solids

The Phases of Water

<table>
<thead>
<tr>
<th>Phase</th>
<th>Temperature (°C)</th>
<th>Density (g/cm³, at 1 atm)</th>
<th>Molar Volume</th>
<th>Molecular View</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gas (steam)</td>
<td>100</td>
<td>$5.90 \times 10^{-4}$</td>
<td>30.5 L</td>
<td></td>
</tr>
<tr>
<td>Liquid (water)</td>
<td>20</td>
<td>0.998</td>
<td>18.0 mL</td>
<td></td>
</tr>
<tr>
<td>Solid (ice)</td>
<td>0</td>
<td>0.917</td>
<td>19.6 mL</td>
<td></td>
</tr>
</tbody>
</table>

Phase Changes

- A solid can be converted into a liquid by heating, and to a gas by heating or reducing the pressure:

Propane (C₃H₈) is stored in tanks at pressures above 2.7 atm, which turns it into a liquid. When the valve is opened, some of the propane evaporates into the gas phase.
Ion-Ion Interactions

- Ionic compounds in general are solids at room temperature, with relatively high melting points (e.g. NaCl = 801°C, KClO₃ = 356°C, MgCl₂ = 714°C, Al₂O₃ = 2072°C, MgO = 2852°C)
- From Coulomb’s Law, the product of the charges $Q₁$ and $Q₂$ in the numerator is always negative— salts are always more stable than the isolated cations and anions.
  - The distance $d$ for ionic compounds is the sum of the ionic radii. Thus, some trends in melting points can be predicted from trends in ion radius.
  - NaBr has a lower melting point (747°C) than NaCl; NaI is even lower (661°C)

$$E = \frac{1}{4\pi\varepsilon₀} \frac{Q₁Q₂}{d}$$  
Coulomb's Law
Intramolecular and Intermolecular Forces

- **Intramolecular forces** operate within each molecule, influencing the chemical properties of the substance (i.e., covalent bonds).
  - These are the forces that hold the atoms in a molecule together. They are very strong forces which result from large charges (on protons and electrons) interacting over very short distances.

- **Intermolecular forces** (van der Waals forces) operate between separate molecules, influencing the physical properties of the substance.
  - These are the forces that hold liquids and solids together, and influence their melting and boiling points. They are weaker forces, because they result from smaller charges, or partial charges, interacting over much larger distances.

  \[ E = \frac{1}{4\pi\varepsilon_0} \frac{Q_1 Q_2}{d} \]

  Coulomb's Law

- To break an O—H single bond in a water molecule, the water must be heated to thousands of degrees C; to completely overcome the intermolecular forces, all you have to do is boil it — 100ºC.

- Intermolecular forces include dispersion forces, dipole-dipole forces, and hydrogen bonds; ion-dipole forces operate between ions and molecules.
Chapter 10 Liquids and Solids

**London Dispersion Forces**

- All atoms and molecules experience **London dispersion forces**, which are attractive forces arising from fluctuations in the electron distribution within atoms or molecules. (Fritz W. London, 1930)

- At any one instant, the random motion of electrons within an atom or molecule may cause the electrons to be clustered more at one end of the particle, giving that end a very small partial negative charge, and creating a temporary instantaneous dipole:

![Diagram of instantaneous dipole](image1.png)

- This instantaneous dipole can create **induced dipoles** in its neighbors by attracting or repelling its electrons:

![Diagram of induced dipoles](image2.png)

- Dispersion forces exist between all molecules, but they are the only forces that exist between nonpolar molecules.
London Dispersion Forces

• Polar molecules and ions can also induce dipoles in nonpolar molecules:

  ![Diagram of dipole-induced dipole interaction](image)

  - This effect partially accounts for the solubility of molecular oxygen (nonpolar) in water and the ability of blood (which contains Fe cations) to bind oxygen.

London Dispersion Forces

• London forces are generally small, with energies in the range of 1-10 kJ/mol (most covalent bonds are well over 100 kJ/mol)

• The exact magnitude of London forces in a molecule depends on the polarizability of the molecule, the ease with which a molecule’s electron cloud can be distorted by a nearby electric field.
  - Polarizability increases with the size of the atom or the molar mass of the molecule, leading to larger London forces and higher boiling points.
  - Molecules with the same molecular weight that are more “spread out” have more surface area, and therefore more London forces between them.
  - Within a family of similar compounds (such as the hydrocarbons), London forces (and therefore boiling point) increase with increasing molar mass.
Chapter 10 Liquids and Solids

**London Dispersion Forces**

<table>
<thead>
<tr>
<th>Noble Gases</th>
<th>MP (°C)</th>
<th>BP (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td>———</td>
<td>-268.8</td>
</tr>
<tr>
<td>Ne</td>
<td>-248.5</td>
<td>-245.9</td>
</tr>
<tr>
<td>Ar</td>
<td>-189.6</td>
<td>-185.8</td>
</tr>
<tr>
<td>Kr</td>
<td>-157.4</td>
<td>-151.7</td>
</tr>
<tr>
<td>Xe</td>
<td>-111.5</td>
<td>-106.6</td>
</tr>
<tr>
<td>Rn</td>
<td>-71.0</td>
<td>-61.7</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Halogens</th>
<th>MP (°C)</th>
<th>BP (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>F₂</td>
<td>-219.7</td>
<td>-188.2</td>
</tr>
<tr>
<td>Cl₂</td>
<td>-101.0</td>
<td>-34.6</td>
</tr>
<tr>
<td>Br₂</td>
<td>-7.3</td>
<td>58.8</td>
</tr>
<tr>
<td>I₂</td>
<td>114.6</td>
<td>184.4</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Hydrocarbons</th>
<th>Compound</th>
<th>Formula</th>
<th>BP (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methane</td>
<td>CH₄</td>
<td>-164.0</td>
<td></td>
</tr>
<tr>
<td>Ethane</td>
<td>C₂H₆</td>
<td>-88.6</td>
<td></td>
</tr>
<tr>
<td>Propane</td>
<td>C₃H₈</td>
<td>-42.1</td>
<td></td>
</tr>
<tr>
<td>Butane</td>
<td>C₄H₁₀</td>
<td>-0.5</td>
<td></td>
</tr>
<tr>
<td>Pentane</td>
<td>C₅H₁₂</td>
<td>36.1</td>
<td></td>
</tr>
<tr>
<td>Hexane</td>
<td>C₆H₁₄</td>
<td>68.9</td>
<td></td>
</tr>
<tr>
<td>Octane</td>
<td>C₈H₁₈</td>
<td>125.5</td>
<td></td>
</tr>
<tr>
<td>Decane</td>
<td>C₁₀H₂₂</td>
<td>174.1</td>
<td></td>
</tr>
<tr>
<td>Eicosane</td>
<td>C₂₀H₄₂</td>
<td>343</td>
<td></td>
</tr>
</tbody>
</table>

![Diagram showing London Dispersion Forces](image)

**London Dispersion Forces**

- All of these molecules are isomers of the formula C₅H₁₂, but the different arrangements of C atoms leads to a different overall shape and surface area:

  - Small contact area, weakest attraction
  - Less surface area, less attraction
  - Large contact area, strong attraction

  ![Small contact area, Less surface area, Large contact area](image)
**Dipole-Dipole Forces**

- **Dipole-Dipole forces** are the attractions between the opposite partial charges in the **permanent dipoles** of polar molecules. The partial charges in one molecule with a permanent dipole are attracted to the opposite partial charges in another polar molecule:

![Dipole-Dipole Forces Diagram]

- Dipole-dipole forces exist between **all** polar molecules (in addition to London forces).

---

**Dipole-Dipole Forces**

- In general, for molecules of the same molecular weight, a polar molecule (dipole-dipole + London) will have a higher boiling point than a nonpolar molecule (London only):

![Butane and Acetone Diagrams]

Butane (C₄H₁₀)
Mol mass = 58 amu
bp = −0.5°C

Acetone (C₃H₆O)
Mol mass = 58 amu
bp = 56.2°C
**Dipole-Dipole Forces**

- Dipole-dipole forces are usually weak, ~3-4 kJ/mol, and are significant only when molecules are in close contact.
- The more polar the molecule, the stronger the dipole-dipole forces and the higher the boiling point.

![Diagram showing dipole-dipole forces](image)

**Hydrogen Bonding**

- **Hydrogen bonding** is an especially strong version of the dipole-dipoles force that occurs in molecules which have H—N, H—O, or H—F bonds:

![Diagram of hydrogen bonding](image)
Hydrogen Bonding

- The electronegativity difference between O, N, and F vs. H is so large that these bonds are especially polar, and the attractions between the opposite partial charges are especially strong.
- H-bonds can have energies up to 40 kJ/mol.

Hydrogen Bonding

- Any molecule that contains an O—H bond, such as ethanol, CH₃CH₂OH, is capable of forming hydrogen bonds:

Ethyl Alcohol
- 46.07 g/mol
- mp -117 °C
- bp 78.5 °C
- density 0.789 g/mol (20°C)
- intoxicant

Dimethyl Ether
- 46.07 g/mol
- mp -139 °C
- bp -25 °C
- density 0.00195 g/mol (20°C)
- refrigerant
Hydrogen Bonding and Boiling Point

- Hydrogen bonding causes water to boil at a much higher temperature than would be “expected” for such a small molecule:

![Graphs showing boiling points of different compounds across periods.](image)

Hydrogen Bonding and Water

- Hydrogen bonding also causes solid water to adopt a more open crystalline structure, which is less dense than the liquid structure (hence, ice floats on water).

![Molecular structures of water and ice.](image)
Hydrogen Bonds and the DNA Double Helix

- **Deoxyribonucleic acid** (DNA), the molecule which encodes genetic information in living cells, consists of a strand of sugar molecules (deoxyribose) connected by phosphate groups; each sugar has one of four bases (A, T, G, C) attached to it. Hydrogen bonds between the bases in the separate strands of DNA cause them to cling together in the famous **double helix** structure:

Ion-Dipole Forces

- **Ion-Dipole forces** are the result of electrical interactions between an ion and the partial charges on a polar molecule.

- These forces are responsible for the ability of polar solvents like water to dissolve ionic compounds.
### Summary: Intermolecular Forces

<table>
<thead>
<tr>
<th>Intermolecular Forces</th>
<th>Formed by the attraction between …</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ion-dipole</td>
<td>an ion and a polar molecule</td>
<td>Na$^+$ and H$_2$O</td>
</tr>
<tr>
<td>Hydrogen bond</td>
<td>molecules which have H on N, O, or F atoms</td>
<td>H$_2$O and H$_2$O; H$_2$O and CH$_3$CH$_2$OH$^-$.</td>
</tr>
<tr>
<td>Dipole - Dipole</td>
<td>two polar molecules</td>
<td>CH$_3$Br and ICl; CH$_3$Br and H$_2$O</td>
</tr>
<tr>
<td>Ion - Induced dipole</td>
<td>an ion and a nonpolar molecule</td>
<td>Fe$^{2+}$ and O$_2$</td>
</tr>
<tr>
<td>Dipole - Induced dipole</td>
<td>a polar molecule and a nonpolar molecule</td>
<td>HCl and Cl$_2$</td>
</tr>
<tr>
<td>London (dispersion) forces</td>
<td>two nonpolar molecules</td>
<td>CH$_4$ and CH$_4$; F$_2$ and F$_2$; CH$_3$ and F$_2$.</td>
</tr>
</tbody>
</table>

### Intermolecular Forces and Solubility

- Polarity also determines whether one liquid mixes with another liquid. The general solubility rule is that “like dissolves like”:
  - Ethanol and water mix because they are both polar. (They are *miscible* with each other — soluble in all proportions.)
  - Oil and gasoline mix with each other because they are both nonpolar.
  - Oil and water don’t mix with each other.
Intermolecular Forces and Solubility

Combinations of Intermolecular Forces

- When larger molecules dissolve in liquid solvents, more than one intermolecular force may be involved. As polar molecules become larger, in general, they become less soluble in water.
Examples: Intermolecular Forces

1. Identify all of the kinds of intermolecular forces in the following substances.
   a. CH₄
   b. CH₃Cl
   c. CH₂Cl₂
   d. CHCl₃
   e. CCl₄
   f. HCl
   g. HF

Examples: Intermolecular Forces

2. What is the strongest kind of intermolecular force in the following substances.
   a. CH₃CH₃
   b. CH₃NH₂
   c. Kr
   d. CH₂O
   e. CO₂
   f. H₂O₂
   g. CH₃OH
Examples: Intermolecular Forces

3. Consider the kinds of intermolecular forces present between the following substances.
   a. CsCl in CH₃Br
   b. CH₃OH in H₂O
   c. CH₃CH₂CH₃ in CCl₄
   d. O₂ and Fe²⁺
   e. CH₃Br in H₂O
   f. HOCH₂CH₂OH in H₂O

Examples: Intermolecular Forces

4. Consider the kinds of intermolecular forces present in the following compounds, and rank the substances in likely order of increasing boiling point.
   H₂S (34 amu)
   C₂H₆ (30 amu, C—C bond length 154 pm)
   CH₃OH (32 amu)
   Ar (40 amu, radius 71 pm)
Examples: Intermolecular Forces

5. Predict which of the following pairs of liquids are miscible with each other.
   a. HOCH₂CH₂OH in H₂O
   b. C₆H₆ (benzene) in pentane (CH₃CH₂CH₂CH₂CH₃)
   c. octanol (CH₃CH₂CH₂CH₂CH₂CH₂CH₂CH₂OH) in octane (CH₃CH₂CH₂CH₂CH₂CH₂CH₂CH₃)

Intermolecular Forces in Everyday Life

- Intermolecular forces such are also responsible for the ability of geckos to climb walls and Post-It notes to stick to surfaces:
Liquids

Intermolecular Forces in Liquids and Solids

- Liquids and solids exist because of the relatively strong intermolecular forces that exist between some compounds.
- Intermolecular forces play a large role in determining the physical properties of liquids.
  - The intermolecular forces between identical molecules of a substance are cohesive forces that hold the molecules of a substance together in the liquid (or solid) phase.
  - The intermolecular forces between different molecules are adhesive forces.
- Solids have relatively stronger intermolecular forces than liquids do.
Surface Tension

- **Surface Tension** — molecules at the surface of a liquid feel a stronger net attraction from the molecules below than do the ones in the interior of the liquid. This causes the surface to act like a skin covering the liquid, and tends to shape the liquid to have the smallest possible surface area (a sphere).
Chapter 10 Liquids and Solids

Surface Tension

- **Surfactants** (soaps, detergents, etc.) disrupt the hydrogen bonding at the surface of water, increasing the ability of the water to “wet” other substances.
- Surface tension decreases as the intermolecular forces weaken.

Viscosity

- **Viscosity** is the measure of a liquid’s resistance to flowing. Viscosity is related to the ease with which individual molecules move around in the liquid and thus to the intermolecular forces present.
  - Viscosity is measured in units of *poise*, P, defined as 1 g cm⁻¹ s⁻¹ (water has a viscosity of 1 cP).
  - Substances composed of small, nonpolar molecules (such as gasoline and benzene) have low viscosities.
  - Polar molecules (such as glycerol), and molecules composed of long chains of atoms (such as oil and grease) have higher viscosities.
  - Extremely concentrated solutions can have a high viscosity (e.g., syrup).
  - The viscosity of a liquid decreases at higher temperatures.
**Capillary Action**

- **Capillarity (capillary action)** — the rise of a liquid inside a thin glass tube, which results from the competition between the attractive forces between the liquid molecules (cohesive forces) and the forces between the liquid molecules and the walls of the glass (adhesive forces).

![Capillary tubes](image)

- For water, the adhesive forces are greater than the cohesive forces, and water rises in the tube, and is raised slightly around the edges, producing an concave meniscus.

- For mercury, the cohesive forces between Hg atoms are greater than the adhesive forces between the Hg atoms and the Si—O bonds in the glass, causing mercury to have a convex meniscus.
The Uniqueness of Water

- Water has a low molecular mass (18.02 g/mol), but is a liquid at room temperature.
  - The highly polar nature of the O—H bond, combined with water’s bent shape, makes water have a significant dipole moment.
  - One water molecule can hydrogen bond with up to four others, resulting in an unusually high boiling point (100°C) for such a small molecule.

The Uniqueness of Water

- The polarity and hydrogen bonding ability of water makes it a great solvent for many substances.
  - Water dissolves many organic substances to some extent (except hydrocarbons).
  - Water dissolves many ionic compounds.
- Water has a high heat capacity, and can absorb a lot of heat without much of a temperature change.
  - Water can therefore be used to regulate temperatures (e.g., in radiators).
  - Water regulates much of the temperature of the Earth.
- Water has a high surface tension, and high capillarity
  - Plants draw water from soil by capillary action.
The Uniqueness of Water

- Water’s hydrogen bonds fix solid water into an open, hexagonal structure, which is less dense than liquid water — hence, ice floats on water.
  - When a lake freezes, the layer of ice that forms at the top of the lake insulates the water below it, often preventing it from freezing solid.
  - Cells are damaged by ice crystals when they are frozen, and often do not survive when they are warmed up again.
  - Flash-freezing can be used to get around this problem, by freezing the water so quickly that ice crystals don’t have a chance to organize into their preferred crystal structure.

Water and Aquatic Life

- When surface water cools in the fall, its density increases to the maximum density of 4°C, causing it to sink, which brings the lower-density, warmer water from the bottom in a continuous cycle. This brings dissolved nutrients and minerals to the surface, where surface-dwelling organisms can use them.
Phase Changes

**Vaporization**

- **Vaporization** — in a container of water, thermal energy keeps the water molecules constantly moving around. Some molecules have more energy than others, and are able to break free at the surface, and go into the vapor phase. The higher the temperature, the greater the average energy of the molecules.

- **Condensation** — gas molecules that slow down can be captured by the liquid.
Vapor Pressure and Dynamic Equilibrium

- When a liquid evaporates in a closed vessel, some liquid molecules acquire enough energy during collisions to evaporate; some of the gas molecules eventually lose energy and return to the liquid.
- When the rate of evaporation equals the rate of condensation, there is a **dynamic equilibrium** between the two phases:

\[
\text{Liquid} \rightleftharpoons \text{Gas}
\]

Vaporization

- The rate of vaporization increases with increasing temperature.
- The rate of vaporization also increases if the surface area increases.
- Molecules with stronger intermolecular forces evaporate with greater difficulty, and are said to be **nonvolatile**.
- Molecules with weaker intermolecular forces evaporate easily, and are said to be **volatile**.
  - Acetone (nail polish remover) and gasoline are more volatile than water. Water is more volatile than motor oil. Motor oil is virtually nonvolatile at room temperature.
The Energetics of Vaporization

• Vaporization is an endothermic process, while condensation is an exothermic process.
  – Sweating causes the body to feel cooler, because as water evaporates from the skin, it absorbs energy from the skin.
  – Steam burns are caused when steam condenses to liquid water on the skin, releasing heat.
• The amount of heat required to vaporize one mole of a liquid to the gas phase is called the heat of vaporization, $\Delta H_{\text{vap}}$.

<table>
<thead>
<tr>
<th>Liquid</th>
<th>Formula</th>
<th>Boiling Point (°C)</th>
<th>$\Delta H_{\text{vap}}$ (kJ/mol) at BP</th>
<th>$\Delta H_{\text{vap}}$ (kJ/mol) at 25°C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>H₂O</td>
<td>100</td>
<td>40.7</td>
<td>44.0</td>
</tr>
<tr>
<td>Isopropyl alcohol</td>
<td>C₃H₈O</td>
<td>82.3</td>
<td>39.9</td>
<td>45.4</td>
</tr>
<tr>
<td>Acetone</td>
<td>C₃H₆O</td>
<td>56.1</td>
<td>29.1</td>
<td>31.0</td>
</tr>
<tr>
<td>Diethyl ether</td>
<td>C₄H₁₀O</td>
<td>34.6</td>
<td>26.5</td>
<td>27.1</td>
</tr>
</tbody>
</table>

Vapor Pressure and Dynamic Equilibrium

• The vapor pressure is the pressure exerted by the vapor at equilibrium with its liquid at that temperature.
  – As the temperature of a liquid increases, the vapor pressure increases.
  – Molecules with weak intermolecular forces have high vapor pressures (i.e., they evaporate easily), and molecules with strong intermolecular forces have low vapor pressures.
Chapter 10 Liquids and Solids

Vapor Pressure and Dynamic Equilibrium

- If a system in dynamic equilibrium is disturbed, the system responds so as to minimize the disturbance and return to a state of equilibrium (Le Châtelier’s Principle).

Boiling Points

- The boiling point of a substance is the temperature at which the vapor pressure equals the external (atmospheric) pressure. (The normal boiling point is the boiling point at 1 atm of pressure.)
  - At the boiling point, molecules in the interior of the liquid also have enough energy to escape the liquid phase, not just those at the surface.
  - The boiling point changes if the applied pressure is changed: water boils at 100°C at sea level, but at 94°C in Denver (elevation 5,280 ft) and 78°C at the top of Mt. Everest (29,035 ft).
  - Pressure cookers use pressure to make water to boil at a higher temperature, allowing food to be cooked faster.
The Clausius-Clapeyron equation

- The relationship between vapor pressure and temperature is given by the Clausius-Clapeyron equation:

\[
\ln P_{\text{vap}} = -\left(\frac{\Delta H_{\text{vap}}}{R}\right)\left(\frac{1}{T}\right) + C \quad \ln \left(\frac{P_2}{P_1}\right) = -\left(\frac{\Delta H_{\text{vap}}}{R}\right)\left(\frac{1}{T_2} - \frac{1}{T_1}\right)
\]

where \( P \) is the vapor pressure, \( R \) is the gas constant in thermodynamic units (8.31 J K\(^{-1}\) mol\(^{-1}\)), \( T \) is the Kelvin temperature, and \( C \) is a constant.

Examples: Intermolecular Forces

1. Calculate the mass of water (in g) that can be vaporized at its boiling point with 155 kJ of heat. \((\Delta H_{\text{vap}} \text{ of H}_2\text{O at 100}^\circ\text{C is 40.7 kJ/mol})\)

Answer: 68.6 g H\(_2\)O
Examples: Vapor Pressure and Boiling Point

2. Which member in each pair of liquids has the higher vapor pressure at a given temperature?

a. CH₃OCH₃ or CH₃CH₂OH

b. C₆H₆ or C₁₀H₈

c. CCl₄ or CBr₄

d. CH₃CH₂CH₂OH or CH₃OH

Examples: Vapor Pressure and Boiling Point

3. Which member of each pair has the higher boiling point?

a. LiCl or HCl

b. C₆H₆ or C₁₀H₈

c. CH₃CH₂CH₂OH or CH₃OH

d. (CH₃)₃N or (CH₃)₂NH
Examples: Clausius-Clapeyron Equation

4. Ether has a vapor pressure of 400. mmHg at 17.9°C and a normal boiling point of 34.6°C. What is the heat of vaporization, $\Delta H_{\text{vap}}$, for ether in kJ/mol?

Answer: 28.5 kJ/mol

Examples: Clausius-Clapeyron Equation

5. The vapor pressure of ethanol at 34.7°C is 100.0 mmHg, and the heat of vaporization of ethanol is 38.6 kJ/mol. What is the vapor pressure of ethanol in mmHg at 65.0°C?

Answer: 388 mmHg
Sublimation and Deposition

From the sublime to the ridiculous there is but one step.
— Napoleon Bonaparte

- The atoms and molecules of a sample in the solid phase are still moving, vibrating around a fixed point.
- Molecules which are vibrating fast enough can escape from the solid and sublime from the solid phase directly into the gas phase. The reverse process is called deposition.
  - Sublimation occurs on ice in freezers, and from frozen foods, leading to freezer burn.
  - Solid carbon dioxide (dry ice) sublimes from the solid phase to the gas phase without melting. Iodine (pictured on right) does this under heating.

Melting (Fusion)

- As heat is added to a solid, the atoms and molecules vibrate faster and faster. At the melting point, the molecules let go of each other enough for the sample to become a liquid. This process is melting or fusion. The opposite process is freezing.
- Once the melting point is reached, the temperature of the sample remains constant until all of the sample has melted.
The Energetics of Melting and Freezing

- Melting is an *endothermic* process, while freezing is an *exothermic* process.
- The amount of heat required to melt one mole of a solid to the liquid phase is called the *heat of fusion*, \( \Delta H_{\text{fus}} \).

### Heats of Fusion of Several Substances

<table>
<thead>
<tr>
<th>Liquid</th>
<th>Formula</th>
<th>Melting Point (°C)</th>
<th>( \Delta H_{\text{fus}} ) (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>H(_2)O</td>
<td>0</td>
<td>40.7</td>
</tr>
<tr>
<td>Isopropyl alcohol</td>
<td>C(_3)H(_8)O</td>
<td>-89.5</td>
<td>5.39</td>
</tr>
<tr>
<td>Acetone</td>
<td>C(_3)H(_6)O</td>
<td>-94.8</td>
<td>5.69</td>
</tr>
<tr>
<td>Diethyl ether</td>
<td>C(<em>2)H(</em>{10})O</td>
<td>-116.3</td>
<td>7.27</td>
</tr>
<tr>
<td>Hydrogen chloride</td>
<td>HCl</td>
<td>-114</td>
<td>1.99</td>
</tr>
<tr>
<td>Carbon tetrachloride</td>
<td>CCl(_4)</td>
<td>-23</td>
<td>2.51</td>
</tr>
<tr>
<td>Sodium chloride</td>
<td>NaCl</td>
<td>801</td>
<td>30.2</td>
</tr>
<tr>
<td>Sodium fluoride</td>
<td>NaF</td>
<td>992</td>
<td>29.3</td>
</tr>
</tbody>
</table>

Heats of Fusion vs. Heats of Vaporization

- \( \Delta H_{\text{fus}} \) is usually less than \( \Delta H_{\text{vap}} \), since transforming a liquid into a gas requires overcoming all intermolecular forces, while they are only partially overcome in going from the solid to the liquid phase.
- By Hess’s Law, \( \Delta H_{\text{subl}} = \Delta H_{\text{fus}} + \Delta H_{\text{vap}} \)
### Phase Changes and Energy Changes

<table>
<thead>
<tr>
<th>Process</th>
<th>$\Delta H^\circ$</th>
<th>State Change</th>
<th>Enthalpy Change</th>
</tr>
</thead>
<tbody>
<tr>
<td>Vaporization</td>
<td>$\Delta H^\circ_{\text{vap}}$</td>
<td>liquid $\rightarrow$ gas</td>
<td>endothermic</td>
</tr>
<tr>
<td>Condensation</td>
<td>$-\Delta H^\circ_{\text{vap}}$</td>
<td>gas $\rightarrow$ liquid</td>
<td>exothermic</td>
</tr>
<tr>
<td>Melting/Fusion</td>
<td>$\Delta H^\circ_{\text{fus}}$</td>
<td>solid $\rightarrow$ liquid</td>
<td>endothermic</td>
</tr>
<tr>
<td>Freezing</td>
<td>$-\Delta H^\circ_{\text{fus}}$</td>
<td>liquid $\rightarrow$ solid</td>
<td>exothermic</td>
</tr>
<tr>
<td>Sublimation</td>
<td>$\Delta H^\circ_{\text{sub}}$</td>
<td>solid $\rightarrow$ gas</td>
<td>endothermic</td>
</tr>
<tr>
<td>Deposition</td>
<td>$-\Delta H^\circ_{\text{sub}}$</td>
<td>gas $\rightarrow$ solid</td>
<td>exothermic</td>
</tr>
</tbody>
</table>

### Examples: Phase Changes

6. Identify the following process, and predict the change in enthalpy, given the following data for the phase changes of water:

- $\Delta H_{\text{fus}} = 6.01$ kJ/mol at 0°C
- $\Delta H_{\text{vap}} = 40.67$ kJ/mol at 100°C

a. $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(s)$

b. $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g)$

c. $\text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(l)$

d. $\text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(s)$
Heating-Cooling Curves

- A **heating-cooling curve** shows the changes that occur when heat is added or removed at constant rate from a sample.
  - *During* a phase change, a change in heat occurs at a constant temperature \( q = n \Delta H_{\text{phase change}} \).
  - *Within* a phase, a change in heat is accompanied by a change in temperature \( q = C_s m \Delta T \).
**Heating Curve for Water**

- **Segment 1:** Solid ice warms from -25°C to 0°C.
  \[ q = C_{s, \text{ice}} m \Delta T \]
- **Segment 2:** Solid ice melts into liquid water at 0°C.
  \[ q = n \Delta H_{\text{fus}} \]
- **Segment 3:** Liquid water warms from 0°C to 100°C.
  \[ q = C_{s, \text{liquid}} m \Delta T \]
- **Segment 4:** Liquid water boils into steam at 100°C.
  \[ q = n \Delta H_{\text{vap}} \]
- **Segment 5:** Steam warms from 100°C to 125°C.
  \[ q = C_{s, \text{steam}} m \Delta T \]

---

**Examples: Heating-Cooling Curves**

7. How much energy (in J) is absorbed when 10.0 g of ice melts to form liquid water? (The heat of fusion for water is 6.01 kJ/mol.)

**Answer:** 3330 J
**Examples: Heating-Cooling Curves**

8. How much energy (in kJ) is needed to heat 270. g of ice from -18.0°C to liquid water at 85.0°C? ($\Delta H_{\text{vap}} = 40.67 \text{ kJ/mol}$, $\Delta H_{\text{fus}} = 6.01 \text{ kJ/mol}$, the specific heat of ice is 2.03 J g$^{-1}$ °C$^{-1}$, the specific heat of liquid water is 4.184 J g$^{-1}$ °C$^{-1}$, and the specific heat of steam is 1.865 J g$^{-1}$ °C$^{-1}$.)

Answer: 196.2 kJ

9. How much heat is required to convert 135 g of ice at -15°C into water vapor at 120°C? ($\Delta H_{\text{vap}} = 40.67 \text{ kJ/mol}$, $\Delta H_{\text{fus}} = 6.01 \text{ kJ/mol}$, the specific heat of ice is 2.03 J g$^{-1}$ °C$^{-1}$, the specific heat of liquid water is 4.184 J g$^{-1}$ °C$^{-1}$, and the specific heat of steam is 1.865 J g$^{-1}$ °C$^{-1}$.)

Answer: 406 kJ
Supercooling and Superheating

• Changes of state do not always occur smoothly and predictably. In the right circumstances, water can be supercooled below 0°C, yet remain liquid, if the water has not reached the level of organization needed to form ice.

• At some point, the correct ordering occurs and the ice rapidly forms, releasing energy in the exothermic process and bringing the temperature back up to the melting point, and the remainder freezes.

Supercooling and Superheating

• A liquid can also be superheated to a temperature above its boiling point, without immediately boiling. This can occur if the liquid is heated rapidly. Once a bubble forms, because its vapor pressure is greater than atmospheric pressure, it can burst before rising to the surface, resulting in “bumping,” and scattering the hot liquid around.
Phase Diagrams

or

Just Another Phase I’m Going Through

Phase Diagrams

• A phase diagram is a chart used to illustrate the stable phases and phase changes of a substance as a function of temperature (on the x-axis) and pressure (on the y-axis).
  – Each region of the phase diagram corresponds to a set of conditions where that particular phase is stable.
  – The lines separating the regions represent the transitions between the different phases, where those two phases are at equilibrium at a particular temperature and pressure.
Navigating Within a Phase Diagram

- Changes in pressure are represented by vertical lines; changes in temperature are represented by horizontal lines.
Parts of a Phase Diagram

- The **triple point** is where all three phase boundaries meet; at this temperature and pressure all three phases are in equilibrium. (For H₂O, this occurs at 0.0098°C and 0.0060 atm.)

- The solid-liquid boundary in water has a slight negative slope because the melting point of ice decreases as the pressure increases.
  - Because ice is less dense than liquid water; increasing the pressure favors the liquid phase.
  - Most substances become less dense when melted, and the solid-liquid boundary line has a positive slope, indicating that the melting point increases with pressure.

Parts of a Phase Diagram

- At the **critical point** the liquid-gas line ends; the **critical temperature**, \( T_c \), is the temperature beyond which a gas cannot be liquefied, no matter how much pressure is applied. (For water, this occurs at 374.4°C and 217.7 atm.)
  - At and above the critical point, the gas and liquid phases become indistinguishable; the substance is said to be a **supercritical fluid**.
  - Supercritical fluids often make good solvents; coffee can be decaffeinated using supercritical carbon dioxide.
Chapter 10 Liquids and Solids

**Under the Sea — Hydrothermal Vents**

*The Blue Planet* (2001)


---

**Phase Diagram for CO₂**

- Solid
- Liquid
- SCF
- Gas
Solids

- Most substances are solids, not liquids or gases.
- Solids can be divided into two broad categories based on the orderliness of their particles:
  - **Crystalline solids** have a well-defined shape and a highly ordered arrangement of particles (e.g., diamond, graphite, ice, quartz).
  - **Amorphous solids** have poorly defined shapes, with randomly arranged particles and no ordered long-range structure (e.g., charcoal, plastics, rubber, glass).
**Chapter 10 Liquids and Solids**

**Types of Crystalline Solids**

- Metals and ionic compounds often form crystalline solids, while large molecules often form amorphous solids, although there are a lot of exceptions to both of these trends.

- Crystalline solids are generally divided into several types, which depend on the type of forces that hold the particles together:
  - Ionic solids
  - Metallic solids
  - Covalent network solids
  - Molecular solids

**Ionic Solids**

- **Ionic solids** consist of ions (monatomic or polyatomic) held together by electrostatic attractions (e.g., NaCl).
  - have generally high melting points
  - don’t conduct electricity except when melted or dissolved in a solvent
  - hard, but brittle (i.e., they shatter instead of bend)
Chapter 10 Liquids and Solids

**Metallic Solids**

- **Metallic solids** are formed by metal atoms, consisting of **metallic bonds** — atomic nuclei held together by a “sea” of delocalized electrons.
  - can be either pure metals or metal alloys
  - conduct electricity and heat well
  - exhibit a metallic luster
  - malleable (can be hammered into sheets) and ductile (can be drawn into wires)
  - can be hard, but because they are malleable, they do not shatter
  - melting points vary widely (Hg is a liquid at room temperature, Ga 29.8°C, Fe 1809°C; W 3422°C)

**Covalent Network Solids**

- **Covalent network solids** are held together by an extended array of covalent bonds (diamond, quartz)
  - usually very hard and strong
  - have high melting points
Molecular Solids

- **Molecular solids** consist of neutral molecules held in place by intermolecular forces (London, dipole-dipole, hydrogen bonds) (e.g. ice, table sugar, iodine).

  - molecular solids made of nonpolar molecules have low melting points.
  - molecular solids made of larger, nonpolar molecules have stronger intermolecular forces, and higher melting points.

Properties of Solids

- Crystalline solids tend to have precise melting points because each interaction between the particles is of the same type, and all have the same strength.

- Amorphous solids tend to soften gradually, rather than melt at a precise temperature, because some of the interactions are of different types, and they do not all have the same strength.

<table>
<thead>
<tr>
<th>Types of Crystalline Solids and Their Properties</th>
</tr>
</thead>
<tbody>
<tr>
<td>Type of Solid</td>
</tr>
<tr>
<td>----------------</td>
</tr>
<tr>
<td>Ionic</td>
</tr>
<tr>
<td>Metallic</td>
</tr>
<tr>
<td>Covalent network</td>
</tr>
<tr>
<td>Molecular</td>
</tr>
</tbody>
</table>
The Solid State: Lattice Structures

Unit Cells

- A **unit cell** in a crystalline solid is its simplest repeating unit.
  - *Lattice points* are where atoms or ions are located.
Unit Cells: Simple Cubic

- The simplest unit cell is the simple cubic unit cell, in which atoms are arranged in a layer directly above atoms in another layer, with the spheres in contact with each other.
  - The spheres are not packed as closely as possible, and fill about 52% of the volume of the unit cell.
  - An atom in this unit cell is directly touching six other atoms (coordination number = 6).

Unit Cells: Simple Cubic

- The edge length of a cubic unit cell is equal to two atomic radii (one atomic diameter).
- At each corner, there is 1/8 of an atom, and since there are 8 corners, a simple cubic unit cell contains a total of 1 atom.
Example: Radius and Density of a Metal

1. The alpha form of polonium crystallizes in a simple cubic unit cell. The edge length of the unit cell is 336 ppm. (a) Determine the radius of a polonium atom. (b) Determine the density of alpha polonium.

Answer: 9.16 g/cm³

Types of Unit Cells

- There are seven different lattice systems, some of which have more than one sub-type, for a total of 14 types of unit cells. For now, we’ll look at the simple cubic unit cell, body-centered cubic unit cell, and face-centered cubic unit cell.
**Unit Cells: Body-Centered Cubic**

- In the **body-centered cubic unit cell** (BCC), there are atoms at the corners, and an atom in the center.
  - the BCC unit cells contain two atoms: \( \frac{1}{8} \) of an atom at the 8 corners, and one in the center.
  - each atom touches 4 atoms in the layer above and 4 in the layer below, for a coordination number of 8.
  - the BCC structure has 68% packing efficiency

![Body-centered cubic structure](image)

**Unit Cells: Face-Centered Cubic**

- In the **face-centered cubic unit cell** (FCC), there are atoms at the corners, and at the center of each face. This is also called **cubic closest packing** (CCP).
  - the FCC unit cells contain four atoms: \( \frac{1}{8} \) of an atom at the 8 corners, and \( \frac{1}{2} \) of an atom at the 6 faces.
  - each atom touches 6 atoms in its own layer, 3 atoms in the layer above and 3 in the layer below, for a coordination number of 12.
  - the BCC structure has 74% packing efficiency

![Face-centered cubic structure](image)
Chapter 10 Liquids and Solids

Cubic and Hexagonal Closest Packing

- There are two types of closest packing in simple metal crystals:
  - In **hexagonal closest packing** (HCP) there are repeating layers of atoms arranged in hexagonal layers, in an ABABAB alternating pattern.
  - In **cubic closest packing** (CCP), there are layers of hexagons in an ABCABC alternating pattern: layer C is not positioned directly above either layer A or B.

The Lattice Systems and Unit Cells

A unit cell is defined by the lengths of its three axes (a, b, and c) and the angles (α, β, and γ) between the axes.
Alloys

- An alloy is produced when a host metal is blended with one or more other elements (metals or nonmetals). This gives the alloy some different properties from the pure host metal.
  - Homogeneous alloys are solid solutions in which the elements added to the host metal are distributed randomly but uniformly.
  - Heterogeneous alloys have matrices of host atoms interspersed with “islands” made of atoms of other elements.

Copper Alloys: Brass and Bronze

- Copper is a relatively soft, easily deformed metal, which reacts with air and water to form hydroxides and carbonates.
- When mixed with zinc (brass) or tin (bronze), it forms alloys which are stronger, and more resistant to reactions with water and air.

The clockwork for the H4 “sea watch” designed by John Harrison (1761)

Bronze Age axes from Germany (3200-600 BC)
Crystal Defects

- Sometimes, defects occur in the repeating pattern of a crystalline solid:
  - **Vacancies** occur when positions that should contain an atom or ion are unoccupied.
  - **Interstitial sites** in between the regular positions of atoms can be occupied by other atoms or molecules.
  - **Substitution impurities** result when a different atom is substituted for the atom which “should” be there; if the atom is too large or too small for that site, it can introduce a distortion into the structure.
  - **Doping** is the process of introducing trace amounts of impurities into a crystal to create defects the produce desirable properties in a material. (This is often done with silicon crystals in order to tune their electronic properties for use in semiconductors or computers.)

Substitutional Alloys

- In a **substitutional alloy**, atoms of the nonhost metal replace host metal atoms in the crystal lattice.
  - These alloys form when the metals have the same crystal lattice shape and atomic radii that are within about 15% of each other.
  - These alloys are often harder because the larger atoms mixed in with the host metal produce “bumps” in the planes of host atoms, making it harder for these planes to slide past each other when a force is applied. They are less malleable, but harder and stronger.
Chapter 10 Liquids and Solids

**Interstitial Alloys**

- In an *interstitial alloy*, atoms with a small enough radius occupy the space between the host metal atoms in its crystal lattice (a “hole” or *interstice*).
- In closest-packed crystal lattices, there are two different sizes of “holes”:
  - *Octahedral holes* are larger — they are surrounded by a cluster of six host atoms in an octahedral shape.
  - *Tetrahedral holes* are smaller — they are surrounded by a cluster of four host atoms.

**Steel**

- Since the radius of C (nonhost) is 77 pm and Fe (host) is 126 pm, a ratio of 0.61 allows C to occupy the octahedral holes in the fcc structure of cooled solid iron (austenite), but not the tetrahedral holes.
- At RT, austenite changes to ferrite, a bcc crystalline form of iron. The octahedral holes are smaller, causing carbon to precipitate out. The clusters of carbon that form inhibit host Fe atoms from sliding past each other, producing a *steel* that is harder and stronger than pure Fe.
- Adding more C to Fe makes a harder steel, but the steel also becomes more brittle.
- By varying the amount of carbon incorporated with the iron, different forms of steel having different levels of malleability or hardness can be produced.
The Solid State: Ionic Crystals

Types of Crystalline Solids: Ionic Solids

• The packing in ionic crystals is more complex than that of metals, because the crystals consist of two or more different kinds of ions that often have different sizes.

• The crystal structures of ionic compounds are a balance between the sizes of the cations and anions, the coordination number of the ions, and the need to maintain charge neutrality.
**Tetrahedral and Octahedral Holes**

- In most ionic compounds, anions are in a closest-packed arrangement. The cations occupy the **interstices** (holes) between the anions:
  - **tetrahedral holes** are found between three anions in one plane and one anion in an adjacent plane
  - **octahedral holes** are found between three anions in one layer and three in the adjacent layer; these are larger than the tetrahedral holes.
- Cations occupy either tetrahedral or octahedral holes, depending on their size.

**Simple Cubic Structure**

- In the case of cesium chloride, CsCl, the cation and anion are about the same size (174 pm and 181 pm), which results in a simple cubic unit cell.
  - There is one Cs\(^+\) in the center and \(\frac{1}{8}\) of 8 Cl\(^-\)s at the corners (or Cl\(^-\) in the center and Cs\(^+\) at the corners).
  - This gives a ratio of 1 Cs\(^+\):1 Cl\(^-\), which is consistent with the formula CsCl.
**The Halite Structure**

- The unit cell of NaCl is a FCC arrangement of Cl\(^{-}\) ions, with the smaller Na\(^{+}\) ions occupying the 12 octahedral holes along the edges of the unit cell and the single octahedral hole in the middle of the cell.
  - There are 4 Cl\(^{-}\) ions in the cell (1/8 in 8 corners, 1/2 in six faces), and 4 Na\(^{+}\) ions (1/4 in 12 edge octahedral holes, and a whole one in the central hole). This gives a ratio of 4 Na\(^{+}\):4 Cl\(^{-}\), which is consistent with the formula NaCl.
  - Each Na\(^{+}\) touches 6 Cl\(^{-}\), and each Cl\(^{-}\) touches 6 Na\(^{+}\).
  - This arrangement of cations and anions is known as the *rock salt structure* (or halite structure).

**The Sphalerite Structure**

- In the mineral sphalerite, ZnS, the S\(^{2-}\) ions (184 pm) form an fcc unit cell, with the smaller Zn\(^{2+}\) ion (74 pm) fitting in the tetrahedral holes.
  - 4 of the 8 tetrahedral holes are occupied by Zn\(^{2+}\) ions.
  - The unit cell contains 4 Zn\(^{2+}\) ions and 4 S\(^{2-}\) ions, which is consistent with the formula ZnS.
  - This pattern of half-filled tetrahedral holes in an fcc unit cell is called the *spalerite structure*. (This is common with Zn, Cd, or Hg with group 16 elements having a 2-charge).
The Fluorite Structure

- In fluorite (CaF₂), there is an fcc arrangement of Ca²⁺ ions at the 8 corners and the middle of the 6 faces, with all 8 tetrahedral holes being filled by the larger F⁻ ions.
  - This has a total of 4 Ca²⁺ ions and 8 F⁻ ions, which is consistent with the formula CaF₂.
  - This structure is known as the fluorite structure. It is also found in SrF₂, CaCl₂, and PbF₂.

- In the anti-fluorite structure (found in Li₂O and K₂S), the smaller cations occupy the tetrahedral holes with the anions in an fcc unit cell.

X-Ray Crystallography

- In order to “visualize” atoms, we must use light that has a shorter wavelength than the atoms themselves.

- In a technique called X-ray diffraction (Max von Laue, 1912; William H. Bragg, William L. Bragg, 1913; Nobel Prizes, 1914 and 1915), X rays are passed through a single crystal of a solid substance; by observing the angles with which the X rays are diffracted and scattered by the crystal, the positions of the atoms can be determined, as well as the lengths of the covalent bonds and the bond angles within the substance.

Bragg equation:
\[ n\lambda = 2d \sin \theta \]
Chapter 10 Liquids and Solids

**X-Ray Crystallography**

![Diagram of X-Ray Crystallography]

**A Crystal Structure**

Photo 51: X-ray diffraction pattern of DNA
Rosalind Franklin (1952)

Courtesy of Cold Spring Harbor Archives. Noncommercial, educational use only.
The Solid State: Other Types of Solids

**Carbon Allotropes**

- Carbon exists in three major allotropes (different molecular forms of the same element).
  - Diamond
  - Graphite
  - Fullerene / buckyballs
**Diamond**

- **Diamonds** consist of an array of C’s linked to four other C’s in a tetrahedral geometry, linked by sp³ orbitals overlapping to form σ bonds.
  - The melting point of diamond (3800°C) is a reflection of the strength of the carbon-carbon covalent bond.
  - Because the bonds are localized σ bonds, diamond is a poor conductor of electricity.
  - Natural diamonds form from graphite under intense heat (>1700K) and pressure (>50,000 atm). Lower-quality synthetic diamonds are made from carbon-rich sources under high temperatures and pressures; these are often used as abrasives or coatings on cutting tools.

**Graphite**

- **Graphite** is a form of carbon in which the C’s are connected in flat sheets, and the atoms are linked by the overlap of sp² orbitals.
  - The side-to-side overlap of unhybridized p orbitals form a network of π bonds that allow graphite to conduct electricity.
  - The individual sheets are held together by relatively weak London forces. The sheets of graphite can therefore slide past each other. This makes graphite soft, and allows it to be used as a lubricant.
  - Graphite is present in soot and smoke, and is used to make lubricants, “lead” pencils, brake linings, steelmaking, batteries, and electrodes.
  - When individual sheets of graphite are isolated, they are referred to as graphene. These sheets behave as semiconductors.
**Fullerene and Nanotubes**

- **Fullerenes** are a form of carbon in which the C’s are grouped in networks of 5- and 6-membered rings arranged in a roughly spherical shape.
  - The first fullerene, C\(_{60}\), was discovered accidentally in 1984; when its discoverers realized what the structure must be, it reminded them of the shape of the geodesic domes designed by the architect Buckminster Fuller, so they named it “buckminsterfullerene.” [Harold Kroto, Robert Curl, and Richard Smalley - 1996 Nobel Prize in Chemistry]
  - Many other fullerenes, or “buckyballs,” were discovered later, including C\(_{20}\), C\(_{70}\), C\(_{80}\), C\(_{82}\), C\(_{84}\), etc.
Fullerene and Nanotubes

- Potential uses of fullerenes:
  - Fullerenes are being investigated for their ability to fit inside the 3D structures of enzymes such as HIV protease.
  - Fullerenes can be modified to transport drug molecules into cells.
  - Fullerenes are also being investigated for their potential use in photodynamic therapy for cancer treatment.
  - Because fullerenes can encapsulate small molecules and ions, they can be used to investigate the composition of ancient atmospheres, or to carry out chemical reactions that can’t take place outside the fullerene cage.

- Carbon nanotubes are fullerenes which are elongated, having diameters of 1-2 nanometers. Nanotubes are extremely strong and lightweight, and are the subject of a great deal of research.

Liquid Crystals

- At some temperatures, some substances exist in a phase which is neither fully liquid nor fully solid. These liquid crystals can move around, as in viscous liquids, but have a restricted range of motion, as in solids.

- The molecules of most liquid crystals have a long, rigid, rod-like shape; when packed together, their molecules tend to orient parallel to each other. The molecules can slide past each other and rotate, but they can’t rotate end over end.

- Two of the most common phases are the nematic phase, in which the ends of the molecules are randomly arranged, and the smectic phase, in which the molecules are arranged in layers.
Liquid Crystals

- The orientation of liquid-crystal molecules is extremely sensitive to small electric fields.
- By applying a small voltage to a small amount of liquid crystal molecules sandwiched between two sheets of polarized glass; the amount of light which is allowed through can be varied, producing a series of pixels which can be used to construct and image.

Liquid Crystal Displays
Conductors, Semiconductors, and Insulators

- **Band theory** models the behavior of electrons in crystalline substances. The molecular orbitals on individual atoms combine to form a collection of orbitals delocalized over the entire crystal.
  - The bonding orbitals combine, decreasing the energy spacings between them until they form a continuous band of energy levels referred to as the **valence band**.
  - The antibonding orbitals combine to form a continuous band at a higher energy called the **conduction band**.
Conductors, Semiconductors, and Insulators

- In metals, the band gap between the valence and conduction bands is so small that electrons are easily promoted to the conduction band. Metals thus act as conductors of electricity and heat.

- In nonmetals, the band gap is too large for electrons to be promoted to the conduction band. Nonmetals thus act as insulators.

- Semiconductors have a small band gap, so electrons can be promoted to the conduction band under some circumstances, resulting in a limited ability to conduct electricity.
  - This ability can be enhanced by the addition of a dopant, which either adds “extra electrons” to the crystal, forcing electrons to go into the conduction band (an n-type semiconductor), or which subtract electrons from the conduction band, creating a positive “hole” that can easily move through the crystal (a p-type semiconductor).

The End