Chapter 12
Liquids, Solids, and Intermolecular Forces

The Physical States of Matter

- Matter can be classified as solid, liquid, or gas based on what properties it exhibits.

<table>
<thead>
<tr>
<th>State</th>
<th>Shape</th>
<th>Volume</th>
<th>Compress</th>
<th>Flow</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solid</td>
<td>Fixed</td>
<td>Fixed</td>
<td>No</td>
<td>No</td>
</tr>
<tr>
<td>Liquid</td>
<td>Indef.</td>
<td>Fixed</td>
<td>No</td>
<td>Yes</td>
</tr>
<tr>
<td>Gas</td>
<td>Indef.</td>
<td>Indef.</td>
<td>Yes</td>
<td>Yes</td>
</tr>
</tbody>
</table>

- Fixed = Keeps shape when placed in a container.
- Indefinite = Takes the shape of the container.
The Structure of Solids, Liquids, and Gases

Solids

- The particles in a solid are packed close together and are fixed in position.
  - Though they are
  - The close packing of the particles results in solids being incompressible.
  - The inability of the particles to move around results in solids retaining their shape and volume when placed in a new container, and prevents the particles from flowing.
Types of Solids

- Crystalline solids

- Amorphous solids

Liquids

- The particles in a liquid are closely packed, but they have some ability to move around.
  - The close packing results in liquids being incompressible.
  - But the ability of the particles to move allows liquids to take the shape of their container and to flow. However, they don’t have enough freedom to escape and expand to fill the container(s).
Gases

- In the gas state, the particles have complete freedom from each other.
  - The particles are constantly flying around, bumping into each other and their container(s).

- In the gas state, there is a lot of empty space between the particles.
  - The particles can be squeezed closer together.

Why Is Sugar a Solid, But Water Is a Liquid?

- The state a material exists in depends on 2 forces
  - The attraction between molecules
  - Their ability to overcome the attraction
### Properties and Attractive Forces

<table>
<thead>
<tr>
<th>Phase</th>
<th>Relative strength of attractive forces</th>
<th>Shape</th>
<th>Volume</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gas</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Liquid</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Solid</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

### Phase Changes: Boiling
- Convert a material in the liquid state into a gas by heating it.
  - Adding heat energy increases the amount of energy of the molecules in the liquid.
  - Overcome the holding them together.
  - Complete freedom to move around and rotate.
Phase Changes: Melting

- Convert a material in the solid state into a liquid by heating it.
  - Adding heat energy increases the amount of the molecules in the solid.
  - Overcome the attractive forces holding them in place.
  - Move around a little and rotate.

Properties of Liquids: Surface Tension

- Liquids tend to
- This tendency causes liquids to have a surface that resists penetration.
- The stronger the attractive force between the molecules,
Properties of Liquids: Surface Tension

- Molecules in the interior of a liquid experience attractions
- Molecules on the surface experience an

Properties of Liquids: Viscosity

- Some liquids flow more easily than others.
- The resistance of a liquid’s flow is called **viscosity**.
  - the attractive
  - the molecule’s shape
Escaping from the Surface

- Evaporation/vaporization
  - The process of molecules of
    - It is a physical change in which a substance is converted from
      - The gaseous form is called
      - It happens at the surface.
        - Molecules on the surface experience a smaller net attractive force than molecules in the interior.
        - All the surface molecules do not escape at once,

- Not all molecules in the sample have the same amount of kinetic energy.

- The average kinetic energy is directly proportional to the
Escaping from the Surface

- Since the higher energy molecules from the liquid are leaving,

- The remaining molecules redistribute their energies, generating more high energy molecules.
- The result is that the liquid continues to evaporate.

Factors Effecting the Rate of Evaporation

- Liquids that evaporate quickly are called volatile liquids, while those that do not are called nonvolatile.

1. Increasing the temperature increases the rate of evaporation.

2. Increasing the surface area increases the rate of evaporation.

3. Weaker attractive forces between the molecules = faster rate of evaporation.
Dynamic Equilibrium: Evaporation and Condensation

Once equilibrium is reached, from that time forward, the amount of vapor in the container will remain the same.

- The vapor pressure of a liquid depends on the

Saturated Vapor Pressure
Boiling

- In an open container, as you heat a liquid the average kinetic energy of the molecules increases, giving more molecules enough energy to escape the surface.
  - So the rate of evaporation increases.
- Eventually, the temperature is high enough for molecules in the interior of the liquid to escape. A phenomenon we call **boiling**.

Boiling Point

- The temperature at which the vapor pressure of the liquid is the same as the atmospheric pressure.
  - The boiling point depends on what the atmospheric pressure is.
    - The temperature of boiling water on the top of a mountain will be cooler than boiling water at sea level.
  - The normal boiling point is the temperature required for the vapor pressure of the liquid to be equal to 1 atm.
Temperature and Boiling

- The amount of heat needed to vaporize one mole of a liquid is called the **heat of vaporization**.
  - $\Delta H_{\text{vap}}$
  - It requires 40.7 kJ of heat to vaporize one mole of water at 100 °C.
  - Always endothermic.
    - Number is +.
- $\Delta H_{\text{vap}}$ depends on
- Since condensation is the opposite process to evaporation, the same amount of energy is transferred but in the opposite direction.

Heat of Vaporization
### Heats of Vaporization of Liquids at Their Boiling Points and at 25 °C

<table>
<thead>
<tr>
<th>Liquid</th>
<th>Chemical formula</th>
<th>Normal boiling point, °C</th>
<th>Δ<em>H</em>\textsubscript{vap} \textit{at} boiling point, (kJ/mol)</th>
<th>Δ<em>H</em>\textsubscript{vap} \textit{at} 25 °C, (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>H\textsubscript{2}O</td>
<td>100</td>
<td>+40.7</td>
<td>+44.0</td>
</tr>
<tr>
<td>Isopropyl alcohol</td>
<td>C\textsubscript{3}H\textsubscript{7}OH</td>
<td>82.3</td>
<td>+39.9</td>
<td>+45.4</td>
</tr>
<tr>
<td>Acetone</td>
<td>C\textsubscript{3}H\textsubscript{6}O</td>
<td>56.1</td>
<td>+29.1</td>
<td>+31.0</td>
</tr>
<tr>
<td>Diethyl ether</td>
<td>C\textsubscript{4}H\textsubscript{10}O</td>
<td>34.5</td>
<td>+26.5</td>
<td>+27.1</td>
</tr>
</tbody>
</table>

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### Example

- Calculate the Mass of Water that Can Be Vaporized with 155 KJ of Heat at 100 °C.
Practice

- How Much Heat Energy, in kJ, is Required to Vaporize 87 g of Acetone, $\text{C}_3\text{H}_6\text{O}$, (MM 58.08) at 25 °C? ($\Delta H_{\text{vap}} = 31.0 \text{ kJ/mol}$)

Temperature and Melting
Heat of Fusion

- The amount of heat needed to melt one mole of a solid is called the **heat of fusion**.
  - \( \Delta H_{\text{fus}} \)
  - Fusion is an old term for heating a substance until it melts, it is not the same as nuclear fusion.
- Since freezing (crystallization) is the opposite process of melting, the amount of energy transferred is the same, but in the opposite direction.
  - \( \Delta H_{\text{crystal}} = -\Delta H_{\text{fus}} \)
- In general, \( \Delta H_{\text{vap}} > \Delta H_{\text{fus}} \) because vaporization requires breaking all attractive forces.

Heats of Fusion of Several Substances

<table>
<thead>
<tr>
<th>Liquid</th>
<th>Chemical formula</th>
<th>Melting point, °C</th>
<th>( \Delta H_{\text{fusion}} ) (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>( \text{H}_2\text{O} )</td>
<td>0.00</td>
<td>6.02</td>
</tr>
<tr>
<td>Isopropyl alcohol</td>
<td>( \text{C}_3\text{H}_7\text{OH} )</td>
<td>-89.5</td>
<td>5.37</td>
</tr>
<tr>
<td>Acetone</td>
<td>( \text{C}_3\text{H}_6\text{O} )</td>
<td>-94.8</td>
<td>5.69</td>
</tr>
<tr>
<td>Diethyl ether</td>
<td>( \text{C}<em>4\text{H}</em>{10}\text{O} )</td>
<td>-116.3</td>
<td>7.27</td>
</tr>
</tbody>
</table>
Example

- Calculate the Mass of Ice that Can Be Melted with 237 kJ of Heat.

Practice

- How Much Heat Energy, in kJ, is Required to Melt 87 g of Acetone, C₃H₆O, (MM 58.08)?
  \( \Delta H_{fus} = 5.69 \text{ kJ/mol} \)
Sublimation

- Sublimation is a physical change in which the solid form changes directly to the gaseous form.
  - Without going through the liquid form.

- Like melting, sublimation is endothermic.

Intermolecular Attractive Forces
Effect of the Strength of Intermolecular Attractions on Properties

- The stronger the intermolecular attractions are, the more energy it takes to separate the molecules.
- Substances with strong intermolecular attractions have higher boiling points, melting points, and heat of vaporization; they also have lower vapor pressures.

Why Are Molecules Attracted to Each Other?

- Intermolecular attractions are a result of attractive forces between opposite charges.
- Larger charge = stronger attraction.
Dispersion Forces

- Also known as London forces or instantaneous dipoles.
  - Caused by distortions in the electron cloud of one molecule inducing distortion in the electron cloud on another.
  - Distortions in the electron cloud lead to a temporary dipole.
    - The temporary dipoles lead to attractions between molecules—dispersion forces.

Strength of the Dispersion Force

- Depends on how easily the electrons can move, or be polarized.
  - The more electrons and the farther they are from the nuclei, the larger the dipole that can be induced.
- Strength of the dispersion force gets larger with larger molecules.
## Dispersion Force and Molar Mass

<table>
<thead>
<tr>
<th>Noble Gas</th>
<th>Molar Mass (g/mol)</th>
<th>Boiling Point (K)</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td>4.00</td>
<td>4.2</td>
</tr>
<tr>
<td>Ne</td>
<td>20.18</td>
<td>27</td>
</tr>
<tr>
<td>Ar</td>
<td>39.95</td>
<td>87</td>
</tr>
<tr>
<td>Kr</td>
<td>83.80</td>
<td>120</td>
</tr>
<tr>
<td>Xe</td>
<td>131.29</td>
<td>165</td>
</tr>
</tbody>
</table>

## Relationship between Dispersion Force and Molecular Size

![Graph showing the relationship between boiling point and period for different types of elements.](image)
Practice—The Following Are All Made of Non–Polar Molecules. Pick the Substance in Each Pair with the Highest Boiling Point.

- CH₄ or C₃H₈
- BF₃ or BCl₃
- CO₂ or CS₂

Permanent Dipoles

- Because of the kinds of atoms that are bonded together and their relative positions in the molecule, some molecules have a **permanent dipole**.
  - Polar molecules.
**Dipole-to-Dipole Attraction**

- Polar molecules have a permanent dipole.
  - A + end and a – end.
- The + end of one molecule will be attracted to the – end of another.

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**Polarity and Dipole-to-Dipole Attraction**

<table>
<thead>
<tr>
<th></th>
<th>Molar Mass (g/mol)</th>
<th>Boiling Point, °C</th>
<th>Dipole size, D</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH₃CH₂CH₃</td>
<td>44</td>
<td>-42</td>
<td>0</td>
</tr>
<tr>
<td>CH₃-O-CH₃</td>
<td>46</td>
<td>-24</td>
<td>1.3</td>
</tr>
<tr>
<td>CH₃ - CH=O</td>
<td>44</td>
<td>20.2</td>
<td>2.7</td>
</tr>
<tr>
<td>CH₃=C≡N</td>
<td>41</td>
<td>81.6</td>
<td>3.9</td>
</tr>
</tbody>
</table>
Attractive Forces

Dispersion forces—All molecules.

Dipole-to-dipole forces—Polar molecules.

Intermolecular Attraction and Properties

- Therefore, the strength of attraction is stronger between polar molecules than between nonpolar molecules of the same size.
Practice—Determine Which of the Following Has Dipole–Dipole Attractive Forces. (EN C= 2.5, F = 4, H = 2.1, S = 2.5)

- $\text{CS}_2$
- $\text{CH}_2\text{F}_2$
- $\text{CF}_4$

Attractive Forces and Properties

- Like dissolves like.
  - Miscible = Liquids that do not separate, no matter what the proportions.

1.

2.

3.
Immiscible Liquids

- When liquid pentane, a nonpolar substance, is mixed with water, a polar substance, the two liquids separate because they are more attracted to their own kind of molecule than to the other.

Hydrogen Bonding

- groups have particularly strong intermolecular attractions.
  - Unusually high melting and boiling points.
  - Unusually high solubility in water.
- This kind of attraction is called a hydrogen bond.
## Properties and H-Bonding

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>MM (g/mol)</th>
<th>Structure</th>
<th>Boiling pt, °C</th>
<th>Melting pt, °C</th>
<th>Solubility in water</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ethane</td>
<td>C₂H₆</td>
<td>30.0</td>
<td><img src="image" alt="Ethane Structure" /></td>
<td>-88</td>
<td>-172</td>
<td>Immiscible</td>
</tr>
<tr>
<td>Ethanol</td>
<td>CH₃OH</td>
<td>32.0</td>
<td><img src="image" alt="Ethanol Structure" /></td>
<td>64.7</td>
<td>-97.8</td>
<td>Miscible</td>
</tr>
</tbody>
</table>

## Intermolecular H-Bonding
Hydrogen Bonding

- When a very electronegative atom is bonded to hydrogen, it strongly pulls the bonding electrons toward it.
- Since hydrogen has no other electrons, when it loses the electrons, the nucleus becomes deshielded.
- The exposed proton acts as a very strong center of positive charge, attracting all the electron clouds from neighboring molecules.

H-Bonds vs. Chemical Bonds

- Hydrogen bonds are not chemical bonds.
- Hydrogen bonds are attractive forces between molecules.
- Chemical bonds are attractive forces that make molecules.
### Relationship between H-Bonding and Intermolecular Attraction

![Graph showing the relationship between boiling point and molar mass for various compounds.](image)

### Attractive Forces and Properties

<table>
<thead>
<tr>
<th>Compound</th>
<th>Molar Mass (g/mol)</th>
<th>Boiling Point, °C</th>
<th>Solubility in water (g/100 g H₂O)</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH₃CH₂OCH₂CH₃</td>
<td>74</td>
<td>34.6</td>
<td>7.5</td>
</tr>
<tr>
<td>CH₃CH₂CH₂CH₂CH₃</td>
<td>72</td>
<td>36</td>
<td>Insoluble</td>
</tr>
<tr>
<td>CH₃CH₂CH₂CH₂OH</td>
<td>74</td>
<td>117</td>
<td>9</td>
</tr>
</tbody>
</table>
Which of the Following Is a Liquid at Room Temperature? (The Other Two Are Gases.)

- formaldehyde, CH₂O
  - 30.03 g/mol
  - polar molecule :: dipole-dipole attractions present
    - polar C=O bond & asymmetric

- fluoromethane, CH₃F
  - 34.03 g/mol
  - polar molecule :: dipole-dipole attractions present
    - polar C–F bond & asymmetric

- hydrogen peroxide, H₂O₂
  - 34.02 g/mol
  - polar molecule :: dipole-dipole attractions present
    - polar H–O bonds & asymmetric
  - H–O bonds :: Hydrogen bonding present

Pick the Compound in Each Pair Expected to Have the Higher Solubility in H₂O.

- CH₃CH₂OCH₂CH₃ or CH₃CH₂CH₂CH₂CH₃.

- CH₃CH₂NHCH₃ or CH₃CH₂CH₂CH₃.

- CH₃CH₂OH or CH₃CH₂CH₂CH₂CH₂OH.
## Types of Intermolecular Forces

<table>
<thead>
<tr>
<th>Type of force</th>
<th>Relative strength</th>
<th>Present in</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Dispersion force</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Dipole–Dipole force</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hydrogen Bond</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Crystalline Solids**
Molecular Crystalline Solids

- Molecular solids are solids whose composite units are molecules.
- Solid held together by intermolecular attractive forces.
- Generally low melting points and $\Delta H_{\text{fusion}}$.

Ionic Crystalline Solids

- Ionic solids are solids whose composite units are formula units.
- Solid held together by attractive forces between cations and anions.
- Generally higher melting points and $\Delta H_{\text{fusion}}$ than molecular solids.
  - Because ionic bonds are stronger than intermolecular forces.
Atomic Crystalline Solids

- Atomic solids are solids whose composite units are individual atoms.
- Solids held together by either covalent bonds, dispersion forces, or metallic bonds.
- Melting points and $\Delta H_{\text{fusion}}$ vary depending on the attractive forces between the atoms.

Classify Each of the Following Crystalline Solids as Molecular, Ionic, or Atomic.

- $\text{H}_2\text{O}(s)$
- $\text{Si}(s)$
- $\text{C}_{12}\text{H}_{22}\text{O}_{11}(s)$
- $\text{CaF}_2(s)$
- $\text{Sc(NO}_3)_3(s)$
Types of Atomic Solids: Covalent

- Covalent atomic solids have their atoms attached by covalent bonds.
  - Effectively, the entire solid is one giant molecule.
- Because covalent bonds are strong, these solids have
- Because covalent bonds are directional, these substances tend to be
- Elements found as covalent atomic solids are C, Si, and B.
- Compounds that occur as covalent atomic solids include SiO$_2$ and SiC.

Types of Atomic Solids: Metallic

- Metallic solids are held together by metallic bonds.
- Metal atoms release some of their electrons to be shared by all the other atoms in the crystal.
- The metallic bond is the attraction of the metal cations for the mobile electrons.
  - Often described as islands of cations in a sea of electrons.
- The luster, malleability, ductility, and electrical and thermal conductivity are all related to the mobility of the electrons in the solid.
- The strength of the metallic bond varies,
Substances with Both Bonding and Nonbonding Attractions

- Some substances have chains or layers of bonded atoms that are then attracted by dispersion forces.
  - Chain substances include grey selenium, polymeric SO₃, and asbestos.
  - Layer substances include graphite, black phosphorus, and mica.

Decide if Each of the Following Atomic Solids Is Covalent, Metallic, or Nonbonding.

- diamond

- neon

- iron
**Water: A Unique and Important Substance**

- Water is found in all three states on Earth.
- As a liquid, it is the most common solvent found in nature.
- Without water, life as we know it could not exist.
  - The search for extraterrestrial life starts with the search for water.

**Water**

- Liquid at room temperature.
  - Most molecular substances that have a molar mass (18.02 g/mol) similar to water’s are gaseous.
- Relatively high boiling point.
- Expands as it freezes.
<table>
<thead>
<tr>
<th>Vocabularies</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting</td>
</tr>
<tr>
<td>Boiling (point)</td>
</tr>
<tr>
<td>Sublimation</td>
</tr>
<tr>
<td>surface tension</td>
</tr>
<tr>
<td>Evaporation/vaporization</td>
</tr>
<tr>
<td>Saturated vapor pressure</td>
</tr>
<tr>
<td>Viscosity</td>
</tr>
<tr>
<td>the heat of vaporization/fusion</td>
</tr>
<tr>
<td>Dispersion Forces/London forces</td>
</tr>
<tr>
<td>Dipole–Dipole force</td>
</tr>
<tr>
<td>Hydrogen Bond</td>
</tr>
<tr>
<td>Solubility</td>
</tr>
</tbody>
</table>