

Chapter 11. Liquids and Intermolecular Forces

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Chapter 11. Liquids and Intermolecular Forces

Common Student Misconceptions

- Students confuse *intermolecular* and *intramolecular* forces.
- Students often do not appreciate how important information from earlier chapters is for the understanding of concepts in this chapter.
- Students have difficulty predicting the relative strength of intermolecular forces involved in different materials.
- Students are unaware that there can be *intramolecular* hydrogen bonding.
- Students confuse cohesion and adhesion.
- Students do not realize that, under the right set of conditions, water also sublimates.
- Students often think that more viscous necessarily means more dense.
- Student often think that the “*liquid*” in liquid crystals refers to these materials being pliable rather than actually being liquid.
- Students often confuse LCD and plasma TV technologies.

Teaching Tips

- Ion-dipole interactions are technically *interparticular* forces.
- The term volatile is often used incorrectly, especially in the media.
- Emphasize that *dynamic equilibrium* is one of the most important concepts used in the latter half of the course.
- The term *boiling point* is often used to refer to the *normal boiling point*, that is, the boiling point at 1.0 atm external pressure. If the external pressure is varied, the boiling point will also vary.
- The term *melting point* is often used when referring to the *normal melting point*, that is, the melting point at 1.0 atm external pressure. Like the boiling point, the melting point also varies with external pressure. The melting point is less sensitive to external pressure than is the boiling point.

11.1 A Molecular Comparison of Gases, Liquids, and Solids^{1,2,3,4,5,6,7,8}

- Physical properties of substances are understood in terms of kinetic-molecular theory:
 - Gases are highly compressible and assume the shape and volume of their container.
 - Gas molecules are far apart and do not interact much with one another.
 - Liquids are almost incompressible and assume the shape but not the volume of the container.
 - Liquids molecules are held together more closely than gas molecules but not so rigidly that the molecules cannot slide past each other.
 - Solids are incompressible and have a definite shape and volume.
 - Solid molecules are packed closely together.
 - The molecules are so rigidly packed that they cannot easily slide past each other.
- Solids and liquids are *condensed phases*.
 - Solids with highly ordered structures are said to be *crystalline*.
- Converting a gas into a liquid or solid requires the molecules to get closer to each other.

¹ “Physical Properties of the Halogens” Movie from Instructor’s Resource CD/DVD

² “Pictorial Analogies I: States of Matter” from Further Readings

³ Figure 11.2 from Transparency Pack

⁴ “Phases of Matter” Activity from Instructor’s Resource CD/DVD

⁵ “Chlorine” 3-D Model from Instructor’s Resource CD/DVD

⁶ “Bromine” 3-D Model from Instructor’s Resource CD/DVD

⁷ “Iodine” 3-D Model from Instructor’s Resource CD/DVD

⁸ “Fluorine” 3-D Model from Instructor’s Resource CD/DVD

- We can accomplish this by cooling or compressing the gas.
- Converting a solid into a liquid or gas requires the molecules to move further apart.
 - We can accomplish this by heating or reducing the pressure on the gas.
- The forces holding solids and liquids together are called intermolecular forces.
- Physical properties of liquids and solids are due to **intermolecular forces**.
 - These are forces *between* molecules.

FORWARDS REFERENCES

- A comparison of phases in terms of entropy will be performed in Chapter 19 (section 19.3).

11.2 Intermolecular Forces^{9,10,11,12}

- The attraction between molecules is an *intermolecular* force.
 - Intermolecular forces are much weaker than ionic or covalent bonds.
- When a substance melts or boils, intermolecular forces are broken.
- When a substance condenses, intermolecular forces are formed.
 - Boiling points reflect intermolecular force strength.
 - A high boiling point indicates strong attractive forces.
- Melting points also reflect the strength of attractive forces.
 - A high melting point indicates strong attractive forces.
- *van der Waals* forces are the intermolecular forces that exist between neutral molecules.
 - These include Dispersion forces, dipole-dipole forces, and hydrogen-bonding forces.
 - Ion-dipole interactions are important in solutions.
 - These are all weak (<15% as strong as a covalent or ionic bond) electrostatic interactions.

Dispersion Forces^{13,14,15,16,17,18,19}

- These are the weakest of all intermolecular forces.
- It is possible for two adjacent neutral molecules to affect each other.
 - The nucleus of one molecule (or atom) attracts the electrons of the adjacent molecule (or atom).
 - For an instant, the electron clouds become distorted.
 - In that instant a dipole is formed (called an *instantaneous* or momentary dipole).
 - One instantaneous dipole can induce another instantaneous dipole in an adjacent molecule (or atom).
 - These two temporary dipoles attract each other.
 - The attraction is called the **dispersion force**, or London dispersion force.
 - Dispersion forces exist between all molecules.
- What affects the strength of a dispersion force?
 - Molecules must be very close together for these attractive forces to occur.
 - **Polarizability** is the ease with which an electron distribution can be deformed.
 - The larger the molecule (the greater the number of electrons) the more polarizable it is.

⁹ “Molecular Handshake; Recognition through Weak Noncovalent Interactions” from Further Readings

¹⁰ “Dancing Crystals: A Dramatic Illustration of Intermolecular Forces” from Live Demonstrations

¹¹ “Hydrogen Chloride” 3-D Model from Instructor’s Resource CD/DVD

¹² “Hydrogen Fluoride” 3-D Model from Instructor’s Resource CD/DVD

¹³ Figure 11.4 from Transparency Pack

¹⁴ Figure 11.6 from Transparency Pack

¹⁵ “Solving the Mystery of Fading Fingerprints with London-Dispersion Forces” from Further Readings

¹⁶ “London-Dispersion Forces and ‘The Wave’” from Further Readings

¹⁷ Figure 11.9 from Transparency Pack

¹⁸ “Pentane” 3-D Model from Instructor’s Resource CD/DVD

¹⁹ “Neopentane (dimethylpropane)” 3-D Model from Instructor’s Resource CD/DVD

- Dispersion forces increase as molecular weight increases.
- Dispersion forces depend on the shape of the molecule.
 - The greater the surface area available for contact, the greater the dispersion forces.
 - Dispersion forces between spherical molecules are smaller than those between more cylindrically shaped molecules.
 - Example: *n*-pentane vs. neopentane.

Dipole-Dipole Forces^{20,21,22,23,24}

- **Dipole-dipole forces** exist between neutral polar molecules.
 - Compare two molecules with similar molecular weight: acetonitrile and propane.
 - With similar molecular weights, we expect similar magnitudes for the dispersion forces.
 - Acetonitrile, (polar) however, has a much higher boiling point than propane (nonpolar)
 - The higher boiling point reflects the dipole-dipole forces present.
- Polar molecules attract each other.
 - The partially positive end of one molecule attracts the partially negative end of another.
- Polar molecules need to be close together to form strong dipole-dipole interactions.
- If two molecules have about the same mass and size, then dipole-dipole forces increase with increasing polarity.

Hydrogen Bonding^{25,26,27,28,29,30,31}

- Experiments show that the boiling points of compounds with H–F, H–O, and H–N bonds are abnormally high.
 - Their intermolecular forces are abnormally strong.
- **Hydrogen bonding** is a special type of intermolecular attraction.
 - This is a special case of dipole-dipole interactions.
 - H-bonding requires:
 - H bonded to a small electronegative element (most important for compounds of F, O, and N).
 - an unshared electron pair on a nearby small electronegative ion or atom (usually F, O, or N on another molecule).
 - Electrons in the H–X bond (X is the more electronegative element) lie much closer to X than H.
 - H has only one electron, so in the H–X bond, the H⁺ presents an almost bare proton to the X[–].
 - Bond energies of hydrogen bonds vary from about 4 kJ/mol to 25 kJ/mol.
 - They are much weaker than ordinary chemical bonds.
- Intermolecular and intramolecular hydrogen bonds have exceedingly important biological significance.
 - They are important in stabilizing protein structure, in DNA structure and function, etc.
- An interesting consequence of H-bonding is that ice floats.

²⁰ “Students as Solids, Liquids, and Gases” from Further Readings

²¹ “Acetonitrile” 3-D Model from Instructor’s Resource CD/DVD

²² “Propane” 3-D Model from Instructor’s Resource CD/DVD

²³ “Dimethyl Ether” 3-D Model from Instructor’s Resource CD/DVD

²⁴ “Acetaldehyde” 3-D Model from Instructor’s Resource CD/DVD

²⁵ Hydrogen Bonding” Animation from Instructor’s Resource CD/DVD

²⁶ “Hydrogen Bonding in Water” Activity from Instructor’s Resource CD/DVD

²⁷ “A People-and-Velcro Model for Hydrogen Bonding” from Further Readings

²⁸ Figure 11.10 from Transparency Pack

²⁹ “Water” 3-D Model from Instructor’s Resource CD/DVD

³⁰ “Ice” 3-D Model from Instructor’s Resource CD/DVD

³¹ “Ammonia” 3-D Model from Instructor’s Resource CD/DVD

- The molecules in solids are usually more closely packed than those in liquids.
 - Therefore, solids are usually more dense than liquids.
- Ice is ordered with an open structure to optimize H-bonding.
 - Water molecules in ice are arranged in an open, regular hexagon.
 - Each δ^+ H points towards a lone pair on O.
 - Therefore, ice is less dense than water.
- Ice floats, so it forms an insulating layer on top of lakes, rivers, etc. Therefore, aquatic life can survive in winter.
- Water expands when it freezes.
 - Frozen water in pipes may cause them to break in cold weather.

Ion-Dipole Forces³²

- An **ion-dipole** force is an interaction between an ion (e.g., Na^+) and the partial charge on the end of a polar molecule/dipole (e.g., water).
- It is especially important for solutions of ionic substances in polar liquids.
 - Example: $\text{NaCl}(aq)$

Comparing Intermolecular Forces^{33,34,35,36,37,38,39,40}

- Dispersion forces are found in all substances.
 - Their strength depends on molecular shapes and molecular weights.
- Dipole-dipole forces add to the effect of dispersion forces.
 - They are found only in polar molecules.
- H-bonding is a special case of dipole-dipole interactions.
 - It is the strongest of the intermolecular forces involving neutral species.
 - H-bonding is most important for H compounds of N, O, and F.
- If ions are involved, ion-dipole (if a dipole is present) and ionic bonding are possible.
 - Ion-dipole interactions are stronger than H-bonds.
- When comparing the relative strengths of intermolecular attractions:
 - If the molecules of the two substances have similar molecular weights and shapes, dispersion forces will be similar in magnitude.
 - Differences in the magnitudes of the intermolecular forces are due to dipole-dipole attractions
 - The greater the molecular polarity, the stronger the intermolecular forces.
 - If the molecules of the two substances have very different molecular weights, dispersion forces will have a significant impact on the intermolecular attractions.
 - Intermolecular attractive forces will generally be greater in the substance with the higher molecular weight.
- Keep in mind that ordinary ionic or covalent bonds are much stronger than these interactions!

FORWARDS REFERENCES

- Soft molecular materials held together by intermolecular forces will be in Chapter 12 (section 12.6).

³² “Put the Body to Them!” from Further Readings

³³ “Intermolecular Forces” Activity from Instructor’s Resource CD/DVD

³⁴ Figure 11.14 from Transparency Pack

³⁵ “Acetic Acid” 3-D Model from Instructor’s Resource CD/DVD

³⁶ “1-Propanol” 3-D Model from Instructor’s Resource CD/DVD

³⁷ “2-Propanol” 3-D Model from Instructor’s Resource CD/DVD

³⁸ “Methylpropane” 3-D Model from Instructor’s Resource CD/DVD

³⁹ “Ethyl Methyl Ether” 3-D Model from Instructor’s Resource CD/DVD

⁴⁰ “Sulfur Dioxide” 3-D Model from Instructor’s Resource CD/DVD

- Breaking of solute-solute and solvent-solvent intermolecular forces and replacing them with solute-solvent interactions will take place in the solution process (Ch. 13).
- The binding between the substrate and the active site in the enzyme action thanks to the intermolecular forces will be discussed in Chapter 14 (section 14.7).
- Hydrogen bonding and the formation of hydrated hydronium ions will be discussed in Chapter 16 (section 16.2).
- Hydrogen bonding will be partially responsible for the relative weakness of HF compared to the strength of other binary acids involving halides (section 16.10).
- Hydrogen bonding and high heat capacity, high melting and boiling points of water will be mentioned again in Chapter 18 (sections 18.2 and 18.3).
- Intermolecular attractions in ice will be discussed in Chapter 19 (section 19.3).
- Hydrogen bonding in alcohols will be discussed in Chapter 24 (section 24.4).
- Hydrogen bonding in the α helix of a protein will be discussed in Chapter 24 (section 24.10).

11.3 Select Properties of Liquids

Viscosity^{41,42}

- **Viscosity** is the resistance of a liquid to flow.
- A liquid flows by sliding molecules over one another.
- Viscosity depends on:
 - the attractive forces between molecules.
 - The stronger the intermolecular forces, the higher the viscosity.
 - the tendency of molecules to become entangled.
 - Viscosity increases as molecules become entangled with one another.
 - the temperature.
 - Viscosity usually decreases with an increase in temperature.

Surface Tension^{43,44,45,46,47,48,49}

- Bulk molecules (those in the liquid) are equally attracted to their neighbors.
- Surface molecules are only attracted inward towards the bulk molecules.
 - Therefore, surface molecules are packed more closely than bulk molecules.
 - This causes the liquid to behave as if it had a “skin”.
- **Surface tension** is the amount of energy required to increase the surface area of a liquid by a unit amount.
- Stronger intermolecular forces cause higher surface tension.
 - Water has a high surface tension (H-bonding)
 - Hg(*l*) has an even higher surface tension (there are very strong metallic bonds between Hg atoms).
- Cohesive and adhesive forces are at play.
 - *Cohesive forces* are intermolecular forces that bind molecules to one another.

⁴¹ “Viscosity Races” from Live Demonstrations

⁴² “1,2,3-Ethanetriol (glycerol)” 3-D Model from Instructor’s Resource CD/DVD

⁴³ “Demonstration of Surface Tension” from Live Demonstrations

⁴⁴ “Polarity, Miscibility, and Surface Tension of Liquids” from Live Demonstrations

⁴⁵ “Lowering the Surface Tension of Water: An Illustration of the Scientific Method” from Live Demonstrations

⁴⁶ “Tears of Wine” from Further Readings

⁴⁷ “Why Do Alcoholic Beverages Have ‘Legs’?” from Further Readings

⁴⁸ “Surface Tension of Water: The Magic Touch” from Live Demonstrations

⁴⁹ “At the Water’s Edge: Surface Spreading and Surface Tension” from Live Demonstrations

- *Adhesive forces* are intermolecular forces that bind molecules to a surface.
- Illustrate this by looking at the meniscus in a tube filled with liquid.
 - The *meniscus* is the shape of the liquid surface.
 - If adhesive forces are greater than cohesive forces, the liquid surface is attracted to its container more than the bulk molecules. Therefore, the meniscus is U-shaped (e.g., water in glass).
 - If cohesive forces are greater than adhesive forces, the meniscus is curved downwards (e.g., Hg(l) in glass)
- **Capillary action** is the rise of liquids up very narrow tubes.
 - The liquid climbs until adhesive and cohesive forces are balanced by gravity.

FORWARDS REFERENCES

- Viscosity of organic compounds, such as polyhydroxyl alcohols, will be mentioned in Chapter 24 (section 24.4).

11.4 Phase Changes⁵⁰

- **Phase changes** are changes of state.
 - Matter in one state is converted into another state.
 - *Sublimation*: solid \rightarrow gas.
 - *Melting or fusion*: solid \rightarrow liquid.
 - *Vaporization*: liquid \rightarrow gas.
 - *Deposition*: gas \rightarrow solid.
 - *Condensation*: gas \rightarrow liquid.
 - *Freezing*: liquid \rightarrow solid.

Energy Changes Accompanying Phase Changes^{51,52,53}

- Energy changes of the system for the above processes are:
 - *melting or fusion*: $\Delta H_{\text{fus}} > 0$ (endothermic).
 - The enthalpy of fusion is known as the **heat of fusion**.
 - *vaporization*: $\Delta H_{\text{vap}} > 0$ (endothermic).
 - The enthalpy of vaporization is known as the **heat of vaporization**.
 - *sublimation*: $\Delta H_{\text{sub}} > 0$ (endothermic).
 - The enthalpy of sublimation is called the **heat of sublimation**.
 - *deposition*: $\Delta H_{\text{dep}} < 0$ (exothermic).
 - *condensation*: $\Delta H_{\text{con}} < 0$ (exothermic).
 - *freezing*: $\Delta H_{\text{fre}} < 0$ (exothermic).
- Generally the heat of fusion (enthalpy of fusion) is less than heat of vaporization.
 - It takes more energy to completely separate molecules, than to partially separate them.
- All phase changes are possible under the right conditions (e.g., water sublimates when snow disappears without forming puddles).
 - The following sequence is endothermic:
heat solid \rightarrow melt \rightarrow heat liquid \rightarrow boil \rightarrow heat gas
 - The following sequence is exothermic:
cool gas \rightarrow condense \rightarrow cool liquid \rightarrow freeze \rightarrow cool solid

⁵⁰ Figure 11.20 from Transparency Pack

⁵¹ “Evaporation As an Endothermic Process” from Live Demonstrations

⁵² “Changes of State” Animation from Instructor’s Resource CD/DVD

⁵³ “Butane” 3-D Model from Instructor’s Resource CD/DVD

Heating Curves^{54,55}

- Plot of temperature change versus heat added is a *heating curve*.
- During a phase change, adding heat causes no temperature change.
 - The added energy is used to break intermolecular bonds rather than cause a temperature change.
 - These points are used to calculate ΔH_{fus} and ΔH_{vap} .
- *Supercooling*: When a liquid is cooled below its freezing point and it still remains a liquid.

Critical Temperature and Pressure⁵⁶

- Gases may be liquefied by increasing the pressure at a suitable temperature.
- **Critical temperature** is the highest temperature at which a substance can exist as a liquid.
- **Critical pressure** is the pressure required for liquefaction at this critical temperature.
 - The greater the intermolecular forces, the easier it is to liquefy a substance.
 - Thus, the higher the critical temperature.
- A substance at temperatures and pressures higher than its critical temperature and pressure is in a state called a **supercritical fluid**.
 - *Supercritical fluid extraction* is used to separate complex mixtures.
 - Supercritical carbon dioxide is widely used for these applications.

FORWARDS REFERENCES

- Supercritical fluids in green chemistry will be discussed in Chapter 18 (section 18.5).
- Thermodynamics of phase changes will be further discussed throughout Chapter 19.

11.5 Vapor Pressure

- Some of the molecules on the surface of a liquid have enough energy to escape the attraction of the bulk liquid.
 - These molecules move into the gas phase.
- As the number of molecules in the gas phase increases, some of the gas phase molecules strike the surface and return to the liquid.
- After some time, the pressure of the gas will be constant.
 - A **dynamic equilibrium** has been established.
 - Dynamic equilibrium is a condition in which two opposing processes occur simultaneously at equal rates.
 - In this case, it is the point when as many molecules escape the surface as strike the surface.
 - **Vapor pressure** of a liquid is the pressure exerted by its vapor when the liquid and vapor are in dynamic equilibrium.
 - The pressure of the vapor at this point is called the equilibrium vapor pressure.

Volatility, Vapor Pressure, and Temperature^{57,58,59}

- If equilibrium is never established, the vapor continues to form.
 - Eventually, the liquid evaporates to dryness.
- Liquids that evaporate easily are said to be **volatile**.
 - The higher the temperature, the higher the average kinetic energy, the faster the liquid evaporates.

⁵⁴ “Heating Curves” Activity from Instructor’s Resource CD/DVD

⁵⁵ Figure 11.22 from Transparency Pack

⁵⁶ “Past, Present, and Possible Future Applications of Supercritical Fluid Extraction Technology” from Further Readings

⁵⁷ “Vapor Pressure vs. Temperature” Animation from Instructor’s Resource CD/DVD

⁵⁸ Figure 11.25 from Transparency Pack

⁵⁹ “Boltzmann Distribution” Activity from Instructor’s Resource CD/DVD

Vapor Pressure and Boiling Point^{60,61}

- Liquids boil when the external pressure at the liquid surface equals the vapor pressure.
 - The **normal boiling point** is the boiling point at 760 mm Hg (1 atm).
- The temperature of the boiling point increases as the external pressure increases.
- Two ways to get a liquid to boil: increase temperature or decrease pressure.
 - Pressure cookers operate at high pressure.
 - At high pressure, the boiling point of water is higher than at 1 atm.
 - Therefore, food is cooked at a higher temperature.

FORWARDS REFERENCES

- Vapor pressure reduction of the solvent in a solution – a colligative property – will be discussed in Chapter 13 (section 13.5).

11.6 Phase Diagrams^{62,63,64,65,66}

- A **phase diagram** is a plot of pressure vs. temperature summarizing all equilibria between phases.
- Phase diagrams tell us which phase will exist at a given temperature and pressure.
- Features of a phase diagram include:
 - *vapor-pressure curve*: generally as temperature increases, vapor pressure increases.
 - *critical point*: critical temperature and pressure for the gas.
 - *supercritical fluid*: state of matter beyond the critical point.
 - *sublimation curve*: separates the solid and gas phases.
 - *melting curve*: separates the solid phase and liquid phases.
 - **normal melting point**: melting point at 1 atm.
 - **triple point**: temperature and pressure at which all three phases are in equilibrium.
 - Any temperature and pressure combination not on a curve represents a single phase.

Phase Diagrams of H₂O and CO₂^{67,68,69,70}

- *Water*:
 - In general, an increase in pressure favors the more compact phase of the material.
 - This is usually the solid.
 - Water is one of the few substances whose solid form is less dense than the liquid form.
 - The melting point curve for water slopes to the left.
 - The triple point occurs at 0.0098 °C and 4.58 mm Hg.
 - The normal melting (freezing) point is 0 °C.
 - The normal boiling point is 100 °C.
 - The critical point is 374 °C and 218 atm.
- *Carbon Dioxide*:
 - The triple point occurs at –56.4 °C and 5.11 atm.

⁶⁰ “Equilibrium Vapor Pressure” Activity from Instructor’s Resource CD/DVD

⁶¹ “The Effect of Pressure on Boiling Point” from Live Demonstrations

⁶² “Journey Around a Phase Diagram” from Further Readings

⁶³ Figure 11.27 from Transparency Pack

⁶⁴ “Phase Diagram” Activity from Instructor’s Resource CD/DVD

⁶⁵ “There is No Perceptible Inflection at the Triple Point” from Further Readings

⁶⁶ “Supercritical Chemistry: Synthesis with a Spanner” from Further Readings

⁶⁷ “A Simple Experiment for Demonstration of the Phase Diagram of Carbon Dioxide” from Live Demonstrations

⁶⁸ Figure 11.28 from Transparency Pack

⁶⁹ “Phase Diagram of Water” Animation from Instructor’s Resource CD/DVD

⁷⁰ “Methane” 3-D Model from Instructor’s Resource CD/DVD

- The normal sublimation point is $-78.5\text{ }^{\circ}\text{C}$. (At 1 atm CO_2 sublimates, it does not melt.)
- The critical point occurs at $31.1\text{ }^{\circ}\text{C}$ and 73 atm.
- Freeze drying: Frozen food is placed in a low pressure (< 4.58 torr) chamber.
 - The ice sublimates.

FORWARDS REFERENCES

- Phase diagrams for a pure solvent and for a solution of a nonvolatile solute will be discussed in Chapter 13 (section 13.5).
- Phase equilibria at melting and boiling points will be further analyzed in Chapter 19.

11.7 Liquid Crystals^{71,72}

- Solids are characterized by their order.
- Liquids are characterized by almost random ordering of molecules.
- There is an intermediate phase where liquids show a limited amount of ordering.
 - **Liquid crystals** are substances that exhibit one or more ordered phases at a temperature above the melting point.
 - Example: The first systematic report of a liquid crystal was cholesteryl benzoate.
 - It melts at $145\text{ }^{\circ}\text{C}$.
 - Between $145\text{ }^{\circ}\text{C}$ and $179\text{ }^{\circ}\text{C}$ cholesteryl benzoate is milky and liquid crystalline.
 - At $179\text{ }^{\circ}\text{C}$ the milky liquid suddenly clears.
 - Cholesteryl benzoate passes through an intermediate liquid crystalline phase.
 - It has some properties of liquids and some of solids.
 - The liquid flows (liquid properties) but has some order (crystal properties).

Types of Liquid-Crystalline Phases^{73,74,75,76,77,78,79}

- Liquid crystal molecules are usually long and rodlike.
 - In normal liquid phases they are randomly oriented.
 - In liquid crystals, the molecules are arranged in specific patterns.
- Classification of liquid crystals depend on the ordering of the molecules.
 - **Nematic liquid crystals** (least ordered).
 - Ordered along the long axis of the molecule only.
 - **Smectic A and smectic C liquid crystals**.
 - Ordered along the long axis of the molecule *and* packed into layers.
 - Intermolecular forces limit the ability of the molecules to slide past one another.
 - **Cholesteric liquid crystals**.
 - Molecules are aligned along their long axis.
 - In addition, the molecules are arranged in layers.
 - Molecules in each plane are twisted slightly relative to molecules in neighboring layers.
 - Many of the molecules are derivatives of cholesterol.
 - An example is cholesteryl octanoate.
 - The molecules are long, flat, and rod-like with a flexible tail.

⁷¹ “Liquid Crystal Inquiries: Add a New Phase to Your Curriculum” from Further Readings

⁷² “Colors in Liquid Crystals” from Live Demonstrations

⁷³ Figure 11.32 from Transparency Pack

⁷⁴ Figure 11.33 from Transparency Pack

⁷⁵ “Liquid Crystals Display New Potential” from Further Readings

⁷⁶ “Preparation and Properties of Cholesteric Liquid Crystals” from Further Readings

⁷⁷ Figure 11.36 from Transparency Pack

⁷⁸ “OLEDs Set to Glow” from Further Readings

⁷⁹ “Shining Examples” from Further Readings

- The flexible tail causes the twist between the layers.
- The flexible tail usually contains many C–C bonds (e.g., the hydrocarbon tail in cholesteryl octanoate).
- The rings in the cholesterol portion are not planar, but give the molecule a flat, sausage-like shape.
- Changes in temperature and pressure cause ordering between layers to change.
- This results in color changes.

Further Readings:

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