Chapter 12

Liquids, Solids, and Intermolecular Forces

Dr. Arno Papazyan

Includes material from the slides provided with Introductory Chemistry, Fifth Edition
Nivaldo J. Tro

By Dr. Sylvia Esjornson
Southwestern Oklahoma State University
Weatherford, OK

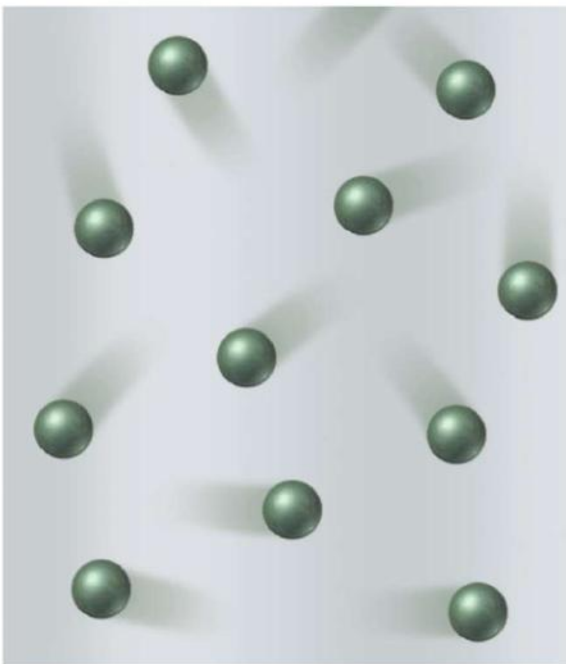
Example of Intermolecular forces in life: Climbing Geckos

- Geckos can adhere to almost any surface
- Recent studies indicate that this amazing ability is related to intermolecular attractive forces
- Geckos have millions of tiny hairs on their feet that branch out and flatten out on the end
  - setae = hairs, spatulae = flat ends
- This structure allows the gecko to have unusually close contact to the surface – allowing the intermolecular attractive forces to act strongly

The state of a sample of matter—solid, liquid, or gas—depends on how strong intermolecular attractive forces are — compared to forces needed to keep the molecules from flying apart from one another

- The molecules and atoms that compose matter are in constant random motion that increases with increasing temperature. The energy associated with this motion is called **thermal energy.**
- The weaker the intermolecular forces (and corresponding energies) relative to thermal energy, the more likely the sample will be gaseous.
- The stronger the intermolecular forces (and corresponding energies) relative to thermal energy, the more likely the sample will be liquid or solid.

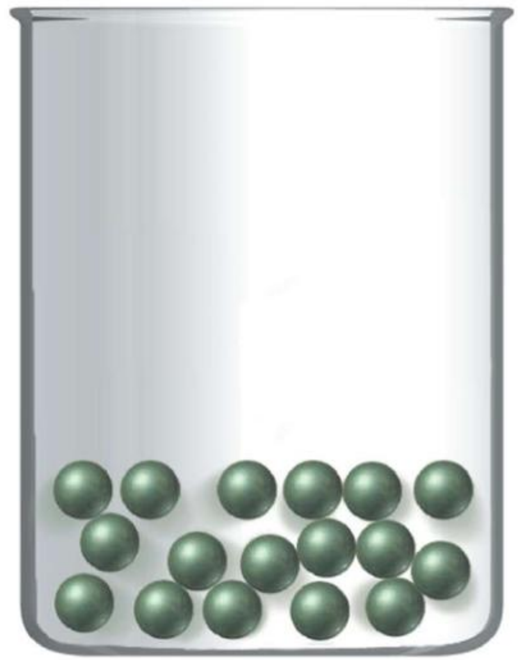
Relative strength of inter-particle forces



Gas

Small

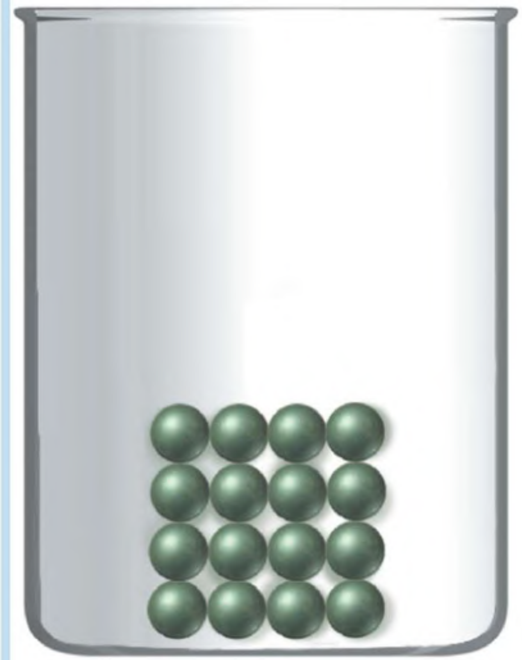
(zero for ideal gases)



Liquid

Large

Enough to keep particles together, but can't pin them down



Solid

Large

Enough to keep particles together & to pin them down

Intermolecular Forces

- Occur between (rather than within) molecules
- For any pair of molecules, much weaker than chemical bonding

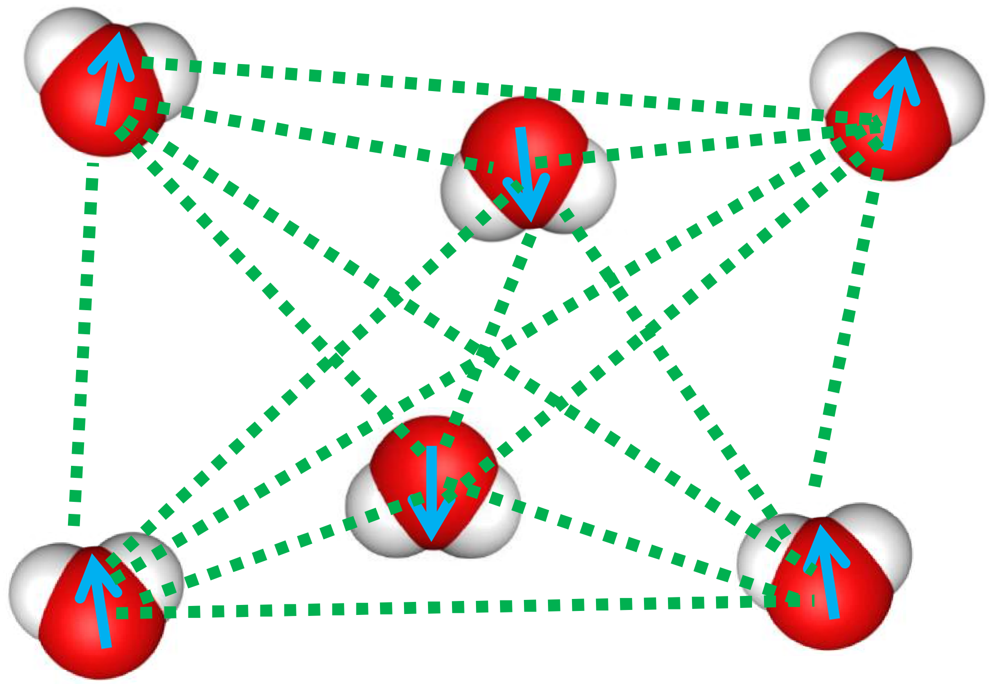
1. Forces created by permanent polarization
   - Between the molecular dipoles: **dipole-dipole**
   - Between N-H, O-H, F-H bonds and N, O, F atoms in molecules: **hydrogen bonding**

2. Forces created by temporary polarization due to fluctuations in electron density: **dispersion (London)**

Forces created by <u>permanent polarization</u>

Case 1: dipole-dipole forces

Between the <u>molecular dipoles</u>: dipole ······ dipole

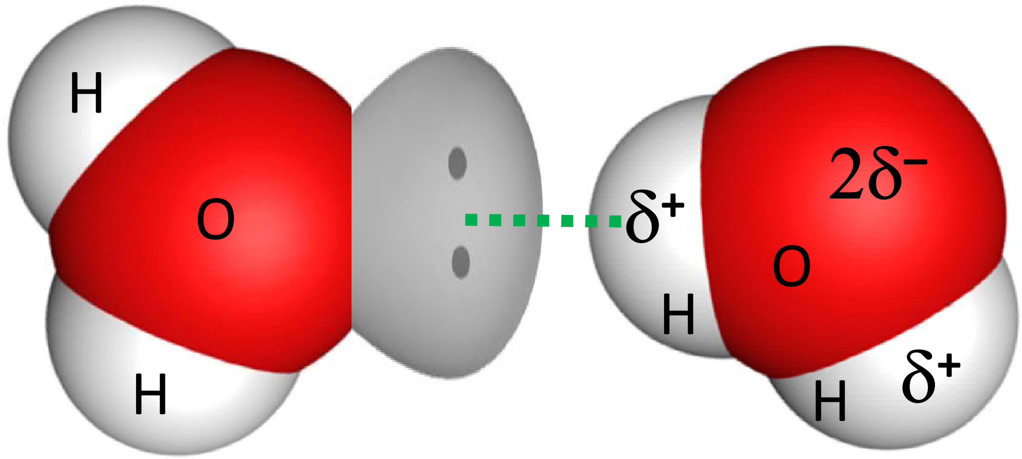


Exists between polar molecules
(molecules with a permanent dipole)

Forces created by <u>permanent polarization</u>

Case 2: “hydrogen bonding” forces

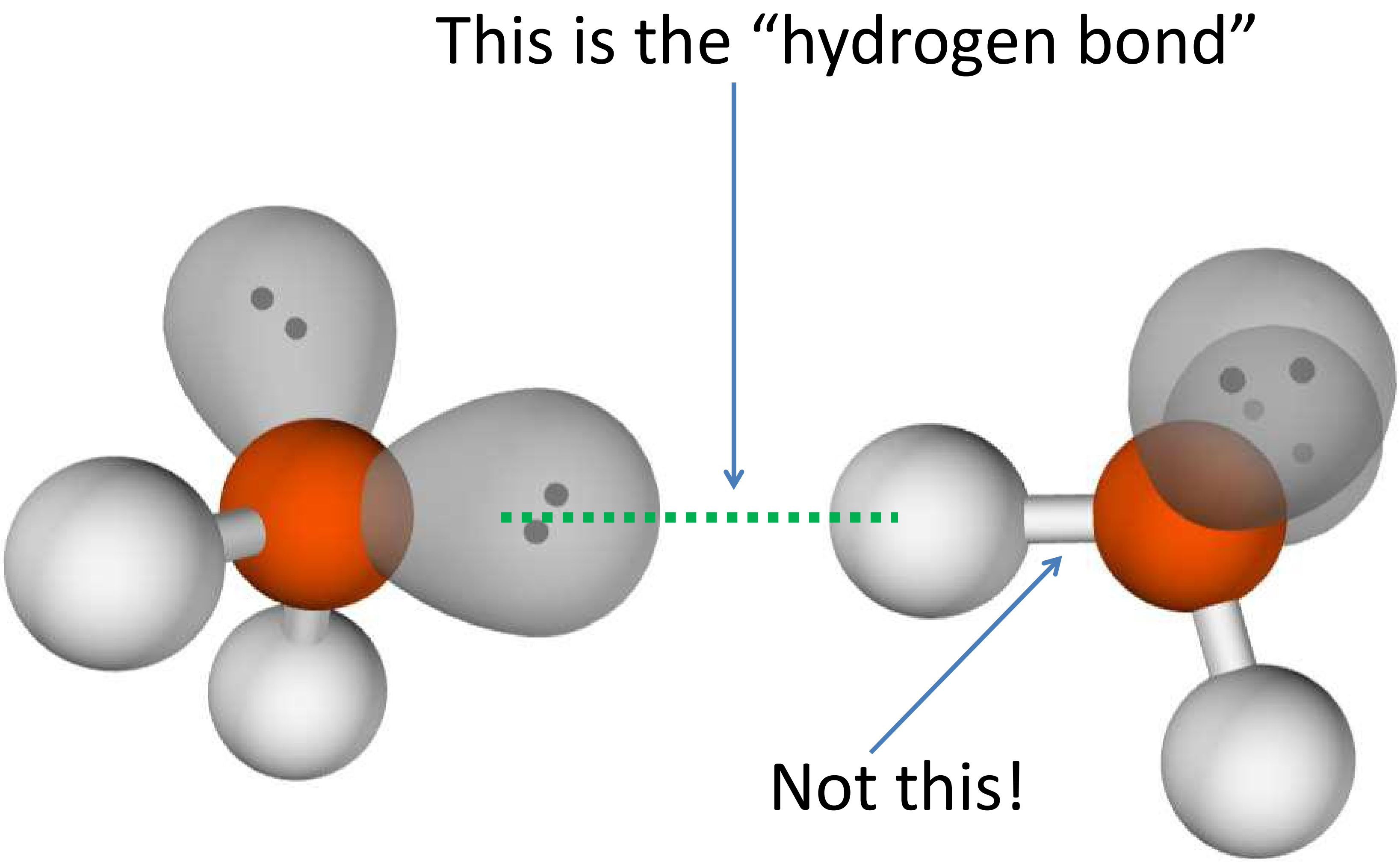
Between N-H, O-H, F-H bonds and N, O, F atoms in molecules



H is “electron-poor” (δ^+) when bonded to N, O, F

H is attracted to the electron-rich N, O, F on <u>another</u> molecule (especially their lone pair electrons)

A “hydrogen bond” has some tiny amount of “chemical bond” character but is <u>**not a chemical bond**</u>


This is the “hydrogen bond”
Not this!

More on Hydrogen Bonding

- It is a strong intermolecular force
- It is an interaction between specific parts of two molecules
- It is <u>not a generic dipole-dipole interaction</u>
- Therefore, it can exist even if the overall molecule is nonpolar
- If we could make a molecule with symmetrically placed bond dipoles that cancel out for the total molecular dipole, it would still have hydrogen bonding interactions with other molecules

London (dispersion) forces

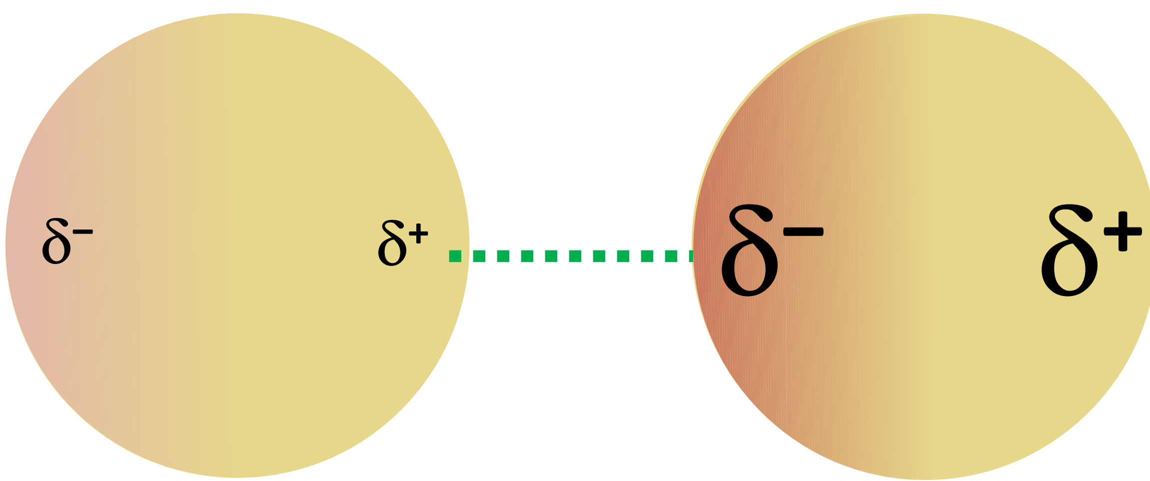
Forces created by <u>temporary polarization due to fluctuations in electron density</u>

All atoms randomly create temporary dipoles (a quantum mechanical effect).

A randomly created dipole on one particle induces a smaller, parallel dipole on nearby particles.

δ^- would push away electrons on other particles, creating a δ^+

Then these dipoles attract, before they pop out of existence.



More on London (dispersion) forces

- All atoms and molecules create this kind of force
  - Because electron clouds always fluctuate
- Dispersion forces exist in all materials
- Dispersion force between any pair of atoms is relatively weak; but can be very strong between molecules with many atoms
- The more electrons there are, the larger the potential fluctuations, and the stronger it gets
- It doesn't take a lot of atoms to give a molecule enough electrons to create stronger dispersion forces than dipole-dipole forces
  - Even for a polar molecule

Molar mass and London (dispersion) forces

- An often-used proxy for the number of electrons on an atom or a molecule is the molar mass.
  - Larger molar mass goes with more electrons per molecule

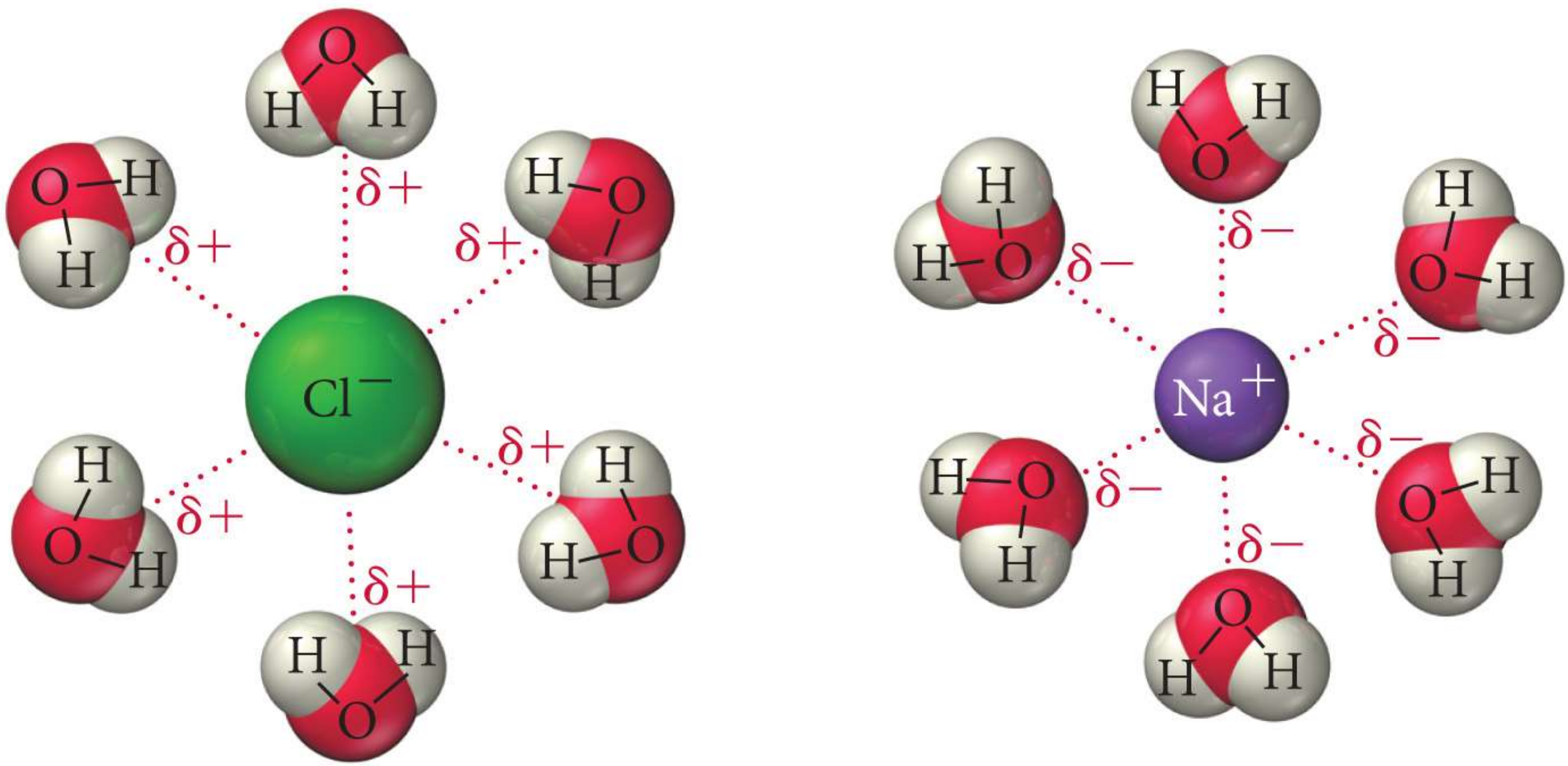
- London forces have nothing to do with how heavy a molecule is. As long as we don’t let that become a misconception, using molar mass as a quick-and-dirty measure of London forces is acceptable.

Ion–Dipole Forces

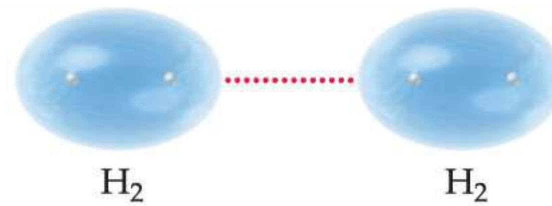
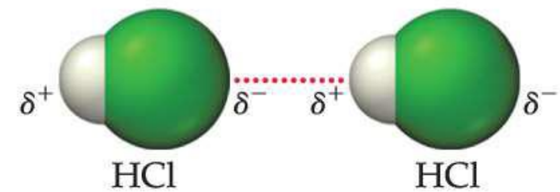
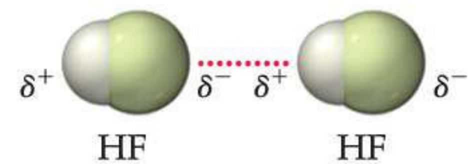
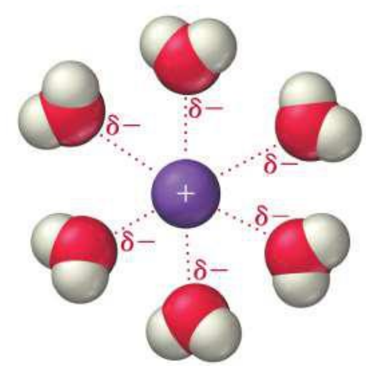
Ion–dipole forces are the strongest intermolecular forces by far

Much larger than dipole-dipole or dispersion (London) forces

For example, ion–dipole forces exist between Na^+ and the negative ends of H_2O molecules and between Cl^- and the positive ends of H_2O molecules.



Different Types of Intermolecular Forces

Type of Force	Relative Strength	Present in	Example
dispersion force (or London force)	weak, but increases with increasing molar mass **Can be the strongest of all, given enough electrons! (it doesn't take huge numbers)**	all atoms and molecules	
dipole–dipole force	moderate	only polar molecules	
hydrogen bond	strong	molecules containing H bonded directly to F, O, or N	
ion–dipole	very strong	mixtures of ionic compounds and polar compounds	

How is the strength of intermolecular forces revealed?

One very common property we use to judge the strength of intermolecular forces:

Boiling point

<u>Stronger</u> intermolecular forces → <u>higher</u> boiling point

Don't let the often-quoted "dispersion forces are weak" statement misguide you

Dispersion forces can easily be stronger than dipole-dipole forces, given enough electrons

HCl	**Cl_2**
Has polar molecules	Has nonpolar molecules
$\Rightarrow$ Has dipole-dipole forces	$\Rightarrow$ No dipole-dipole forces
Has dispersion forces (from 1 + 17 = 18 e^-)	Has only dispersion forces (from 17 + 17 = 34 e^-)
Boiling point = 188 K	**Boiling point = 238 K**

Stronger intermolecular forces → higher boiling point

Cl_2 has stronger intermolecular forces than the polar HCl molecule, even if it only has dispersion forces.

But hydrogen-bonding forces are harder to compete with

Example:

Predict which molecule has stronger intermolecular forces, and explain:

N_2 H_2O

- H_2O has hydrogen bonding (in addition to some London dispersion).
- N_2 exhibits London dispersion forces only, and with more electrons than H_2O, they are stronger than for H_2O, but it would take many more electrons to compete with H-bonding between H_2O molecules.

Predict which gas would behave more ideally at the same pressure and temperature, and explain.

HCN or N_2

Both molecules have the same number of electrons.

So their London forces are very similar.

HCN is a polar molecule, so it has dipole-dipole forces.

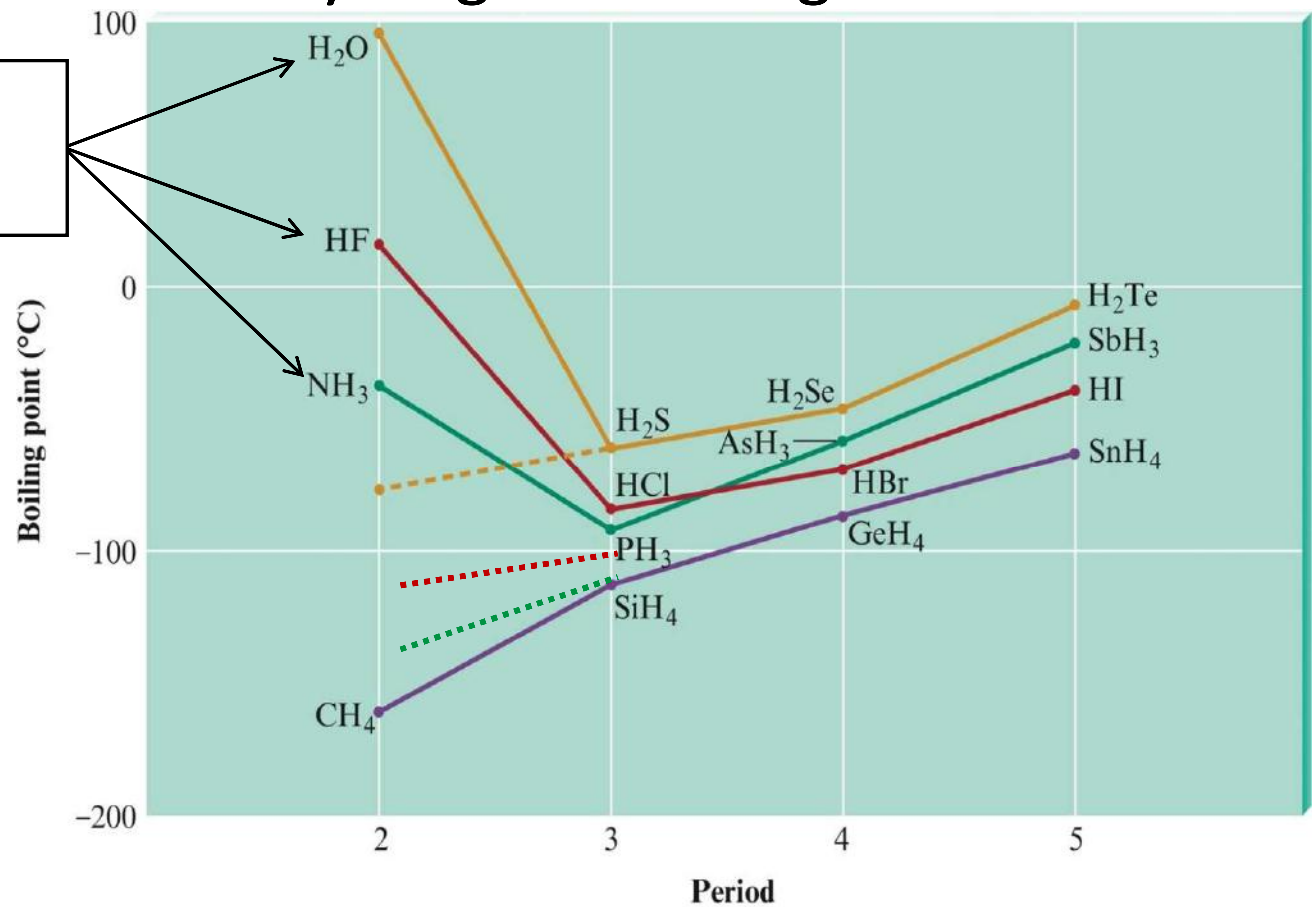
N_2 is nonpolar (lacks a dipole), so it has no dipole-dipole forces.

Since N_2 has weaker intermolecular forces, it would behave more like an ideal gas.

Boiling point, an excellent measure of intermolecular forces, shows, for example, when hydrogen bonding happens

Hydrogen compounds of **F**, **O**, and **N** have abnormally high boiling points due to their hydrogen bonding

They defy the trend in their Group



Some other properties that get larger with stronger intermolecular forces:

- Enthalpy of vaporization
  - Energy required to vaporize molecules into gas phase
- Melting point
  - Less reliable; molecules with unwieldy shapes don't fit into a crystal lattice and have lower melting points
- Viscosity
  - Resistance to flow
  - Also higher for molecules that can get "tangled up"
- Surface tension
  - How hard it is to create more liquid surface

Intermolecular Forces in Action: Surface Tension and Viscosity

- The most important manifestation of intermolecular forces is the very existence of liquids and solids.
- Without intermolecular forces, solids and liquids would not exist and all matter would be gaseous.
- In liquids, we can observe several other manifestations of intermolecular forces, including surface tension and viscosity.

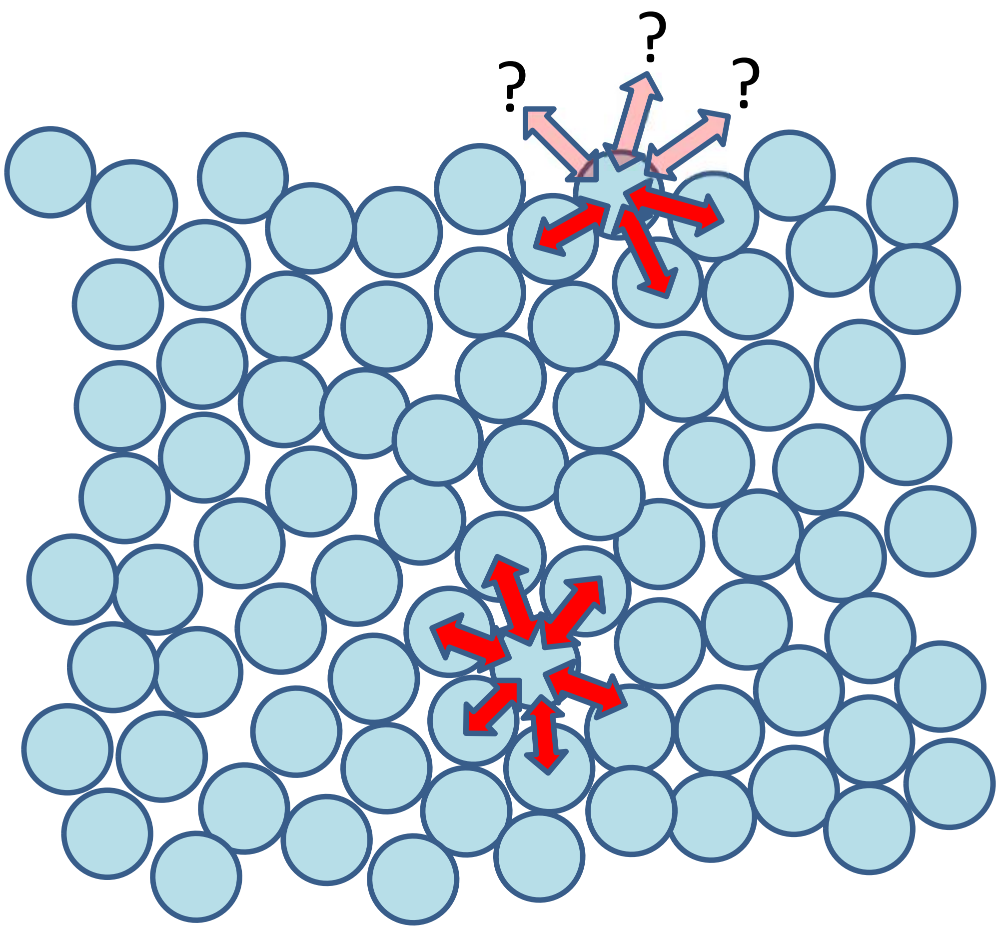
Surface Tension

- A paper clip will float on water if it is carefully placed on the surface of the water. It is held up by surface tension.
- You can't float a paper clip on gasoline because the intermolecular forces among the molecules composing gasoline are weaker than the intermolecular forces among water molecules.

Dr. P's note:
The liquid also should not "wet" the object.

Surface Tension

Molecules at the surface have fewer neighbors to stabilize them (so they are at higher energy)



To minimize the number of molecules with missing neighbors, the "surface tension" force minimizes the surface area of the liquid.

Water droplet in the space station

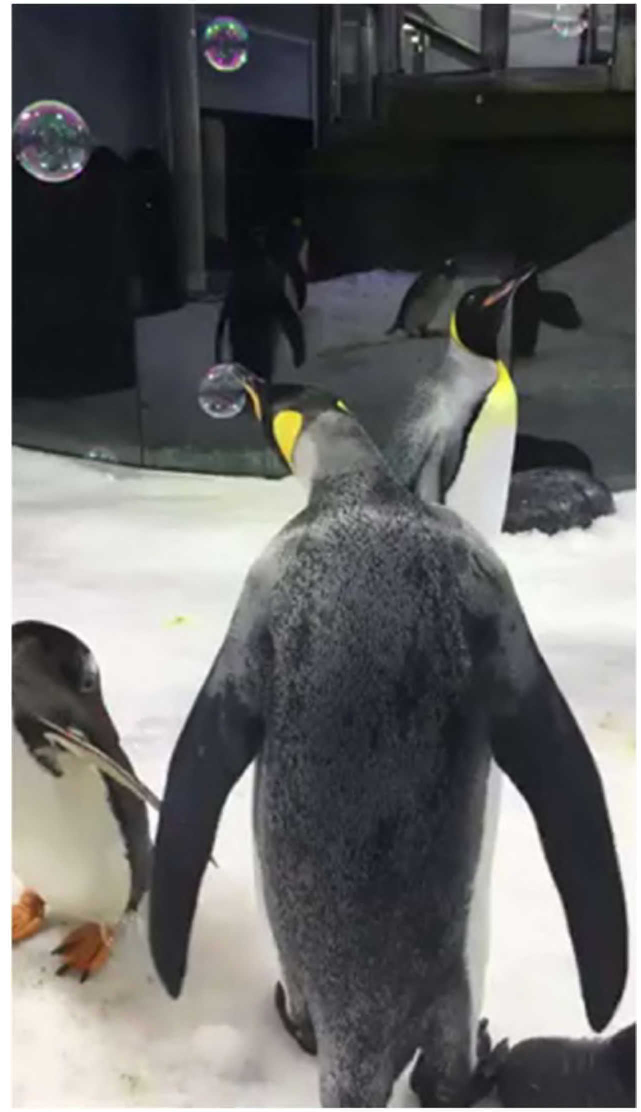
A sphere has the minimum area for a given volume

<u>Surface tension favors a spherical shape</u>

If there are no other forces, droplets are spherical

Bubbles also tend to be spherical unless distorted by other forces (like gravity)

For the same reason.



Viscosity

- **Viscosity** is the resistance of a liquid to flow.
- Liquids that are viscous flow more slowly than liquids that are not viscous.
- Motor oil is more viscous than gasoline.
- Maple syrup is more viscous than water.
- Viscosity is greater in substances with stronger intermolecular forces because molecules cannot move around each other as freely, hindering flow.
- Long molecules, such as the hydrocarbons in motor oil, tend to form viscous liquids because of molecular entanglement.
  - But remember: long molecules also have strong dispersion forces that contribute to viscosity as well

Viscosity

- Maple syrup is more viscous than water because its molecules interact strongly, so they cannot flow past one another easily.

Evaporation, Condensation, and Thermal Energy

The rate of vaporization increases with the following:

- Increasing surface area
- Increasing temperature
- Decreasing strength of intermolecular forces
- Liquids that evaporate easily are termed **volatile,** while those that do not vaporize easily are termed **nonvolatile.**

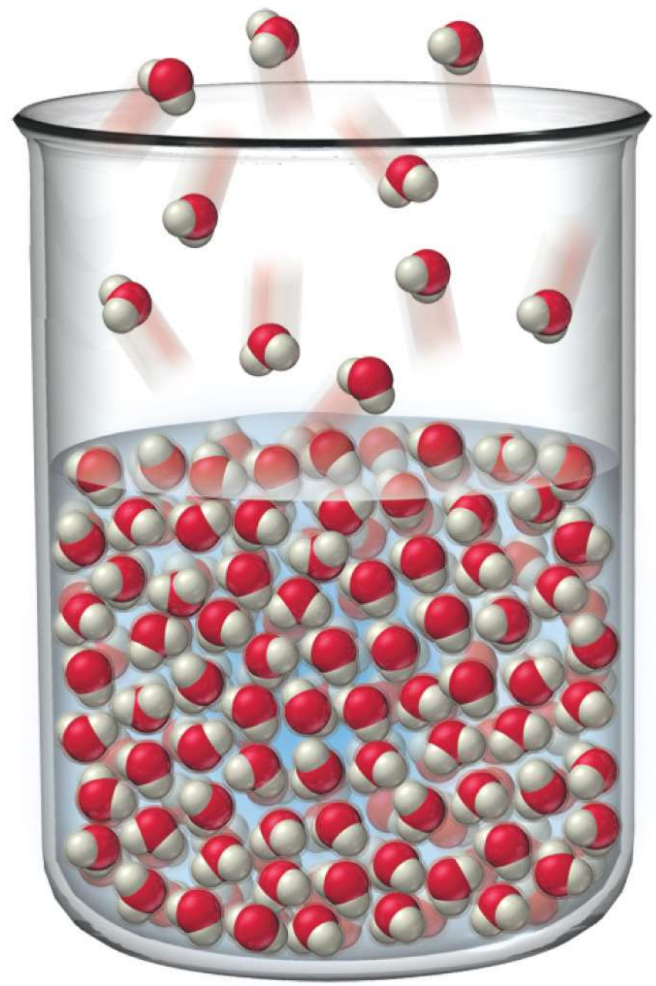
Evaporation

Not all molecules in a liquid have the same kinetic energy.

They constantly bump around, exchange energy, and some molecules end up with more kinetic energy than others.

The most energetic molecules can break away into the gas state in the process called *evaporation.*

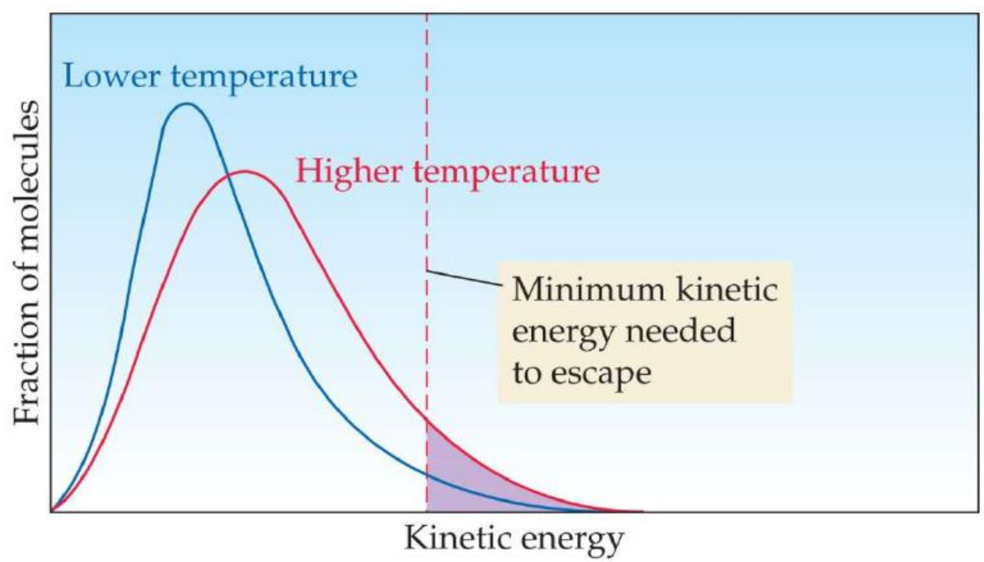
In evaporation or vaporization, a substance is converted from its liquid state into its gaseous state.



Evaporation, Condensation, and Thermal Energy

- At a given temperature, a sample of molecules or atoms will have a distribution of kinetic energies.
- Only a small fraction of molecules has enough energy to escape.

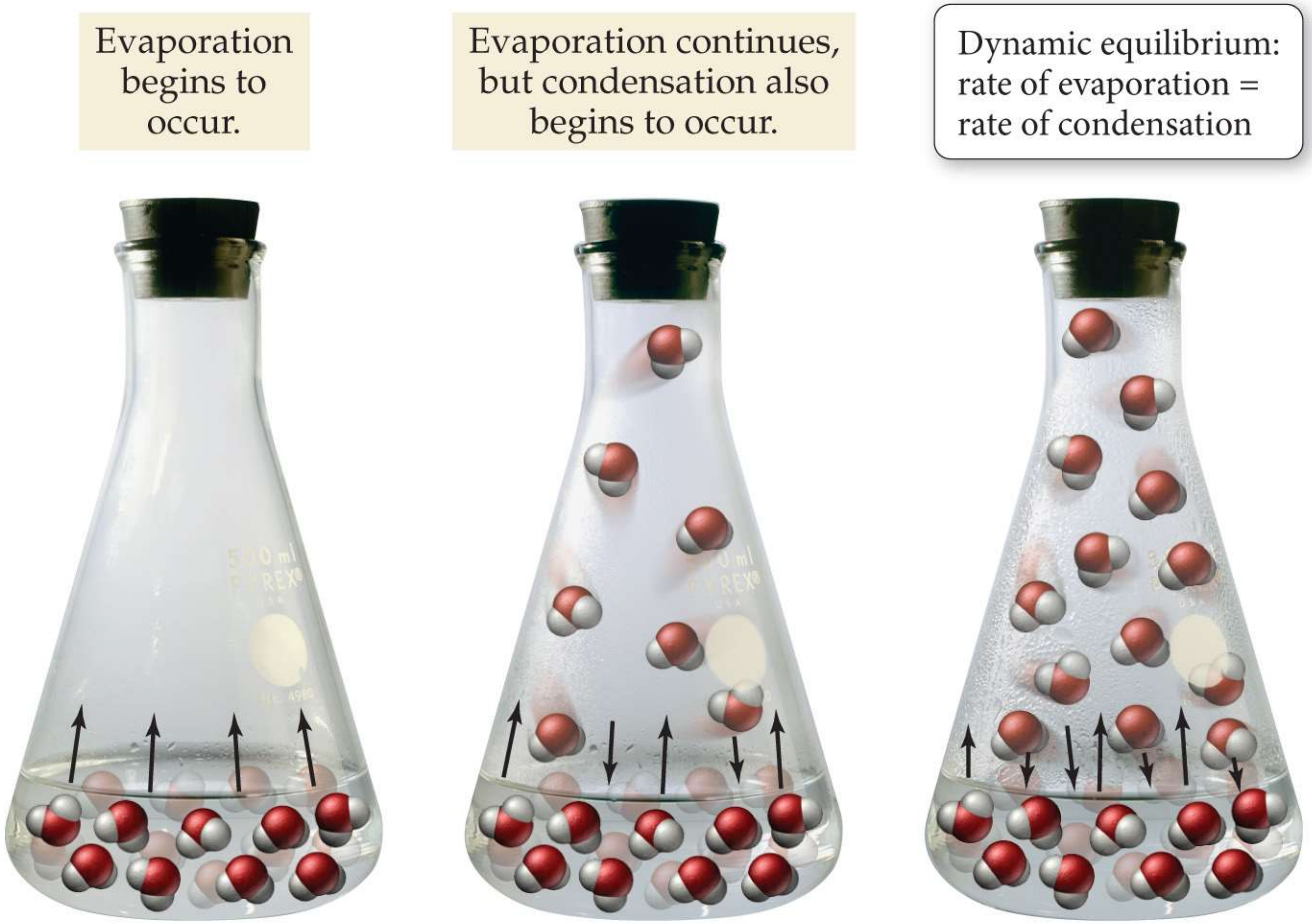
At a higher temperature, the fraction of molecules with enough energy to escape increases.



Evaporation and Condensation

- **Condensation** is a physical change in which a substance is converted from its gaseous state to its liquid state.
- Evaporation and condensation are opposites: Evaporation is a liquid turning into a gas, and condensation is a gas turning into a liquid.
- At the point where the rates of condensation and evaporation become equal, **dynamic equilibrium** is reached and the number of gaseous water molecules above the liquid remains constant.
- The **vapor pressure** of a liquid is the partial pressure of its vapor in dynamic equilibrium with its liquid.

Evaporation and Condensation



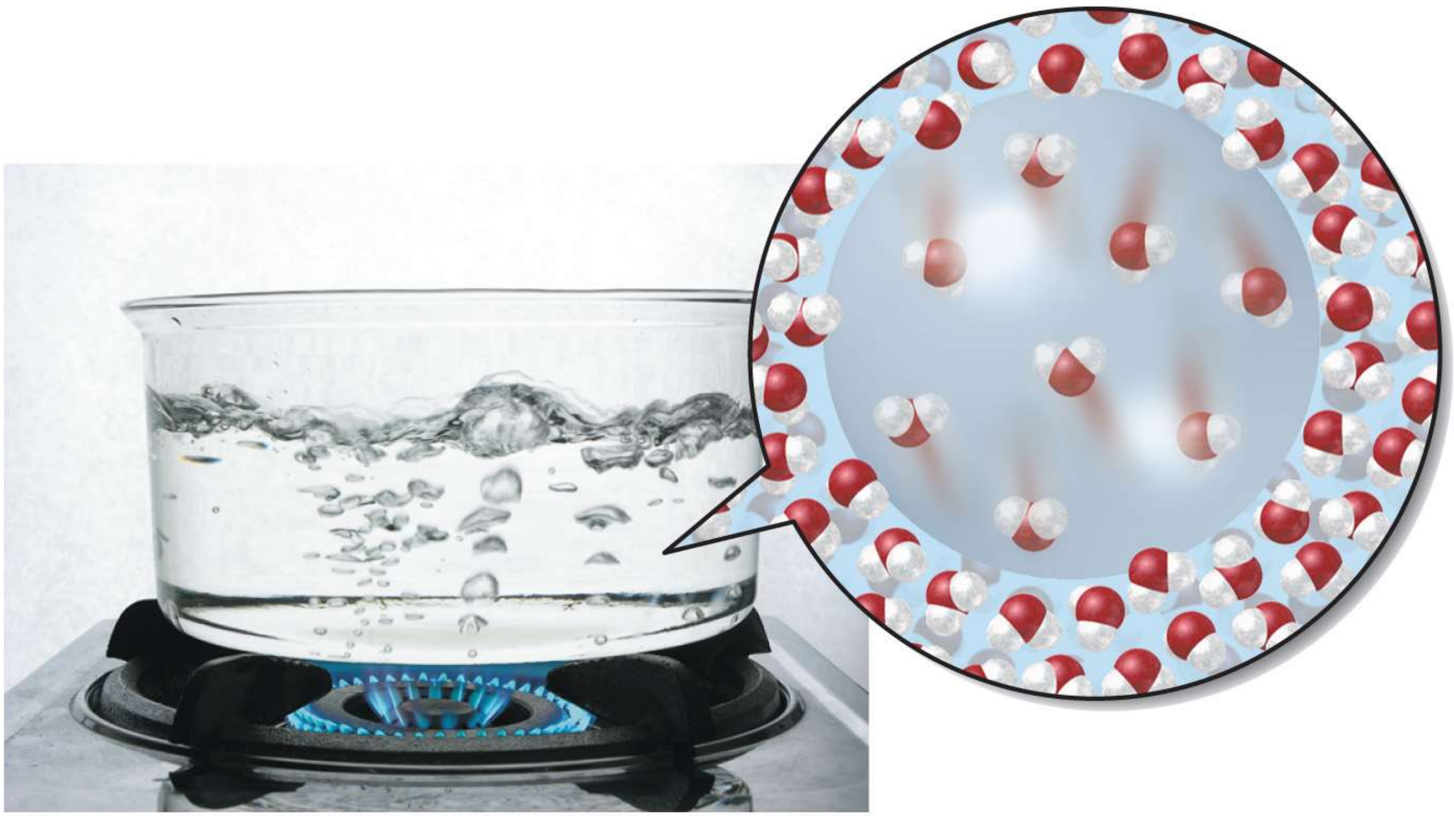
Vapor Pressure

Vapor pressure increases with the following:

- Increasing temperature
- Decreasing strength of intermolecular forces
  - Can’t change it for a given substance at a given T
  - Only useful for comparing different substances

Vapor pressure is independent of surface area because an increase in surface area at equilibrium equally affects the rate of evaporation and the rate of condensation.

Boiling During boiling, thermal energy is enough to cause water molecules in the interior of the liquid to become gaseous, forming bubbles containing gaseous water molecules.

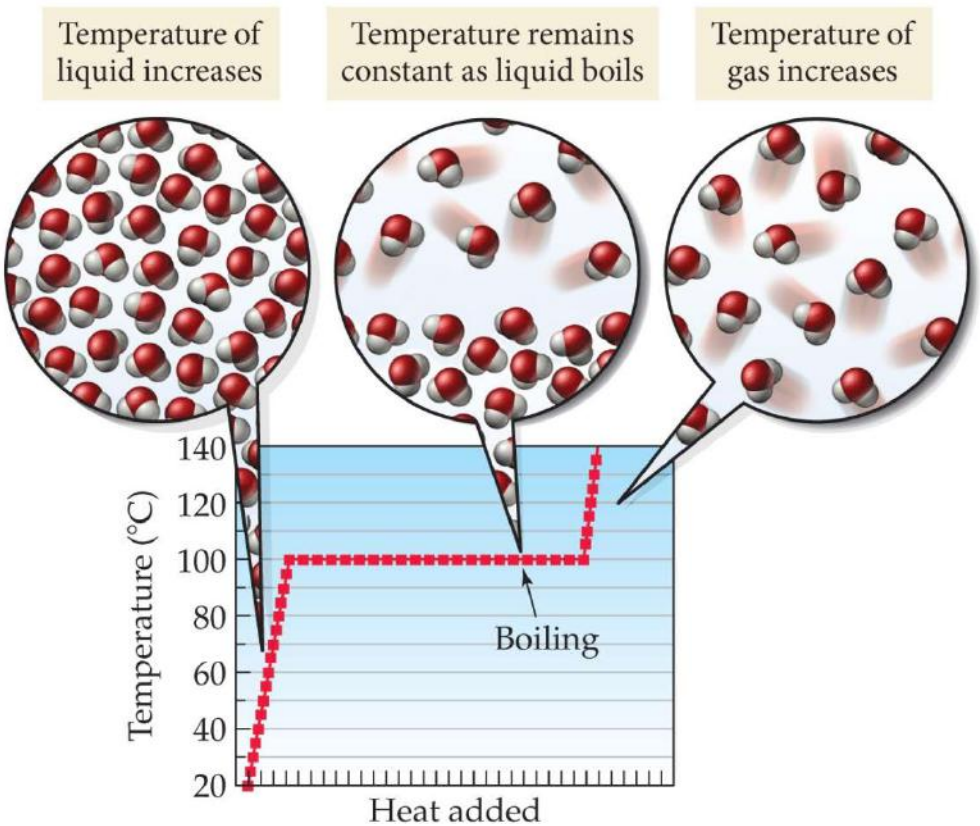


Nature of phase transitions for pure substances: They occur at a constant temperature (at a given pressure), characteristic to the pure substance

- Once the boiling point of a liquid is reached, additional heating only causes more rapid boiling; it does not raise the temperature of the liquid above its boiling point.
- A mixture of boiling water *and* steam will always have a temperature of 100 °C (at 1 atm pressure).
- Only after all the water has been converted to steam can the temperature of the steam rise beyond 100 °C.

A “Heating Curve” for Water

The temperature of water as it is heated from room temperature to the boiling point.



During boiling, the temperature remains at 100 °C until all the liquid is evaporated.

Energetics of Evaporation

- Evaporation is *endothermic*—when a liquid is converted into a gas, it absorbs heat because energy is required to break molecules away from the rest of the liquid.
- Our bodies use the endothermic nature of evaporation for cooling.
- When we overheat, we sweat, causing our skin to be covered with liquid water.
- As this water evaporates, it absorbs heat from our bodies, cooling us down.
- High humidity slows down evaporation, preventing cooling. When the air already contains high amounts of water vapor, sweat does not evaporate as easily, making our cooling system less efficient.

Energetics of Condensation

- Condensation, the opposite of evaporation, is *exothermic* — heat is released when a gas condenses to a liquid.
- As steam condenses to a liquid on your skin, it releases heat, causing a severe burn.
- The exothermic nature of condensation is also the reason that winter overnight temperatures in coastal cities, which tend to have water vapor in the air, do not get as low as those in deserts, which tend to have dry air.
- As the air temperature in a coastal city drops, water condenses out of the air, releasing heat and preventing the temperature from dropping further.
- In deserts, there is little moisture in the air to condense, so the temperature drop is greater.

Heat of Vaporization

- The amount of heat required to vaporize 1 mol of liquid is the **heat of vaporization (ΔH_{vap})**.
- The heat of vaporization of water at its normal boiling point (100 °C) is 40.7 kJ/mol.
- ΔH_{vap} is positive because vaporization is endothermic; energy must be added to the water to vaporize it.
- The same amount of heat is involved when 1 mol of gas condenses, but *the heat is emitted* rather than absorbed.
- ΔH_{cond} is negative because condensation is exothermic; energy is given off as the water condenses.

Heat of Vaporization of Water

$$H_2O(l) \rightarrow H_2O(g) \qquad \Delta H = +40.7 \text{ kJ (at 100 °C)}$$

$$H_2O(g) \rightarrow H_2O(l) \qquad \Delta H = -40.7 \text{ kJ (at 100 °C)}$$

Heats of Vaporization

TABLE 12.2 Heats of Vaporization of Several Liquids at Their Boiling Points and at 25 °C

Liquid	Chemical Formula	Normal Boiling Point (°C)	Heat of Vaporization (kJ/mol) at Boiling Point	Heat of Vaporization (kJ/mol) at 25 °C
water	H_2O	100.0	40.7	44.0
isopropyl alcohol (rubbing alcohol)	C_3H_8O	82.3	39.9	45.4
acetone	C_3H_6O	56.1	29.1	31.0
diethyl ether	$C_4H_{10}O$	34.5	26.5	27.1

- Use the heat of vaporization of a liquid to calculate the amount of heat energy required to vaporize a given amount of that liquid.
- Use the heat of vaporization as a conversion factor between moles of the liquid and the amount of heat required to vaporize it.

EXAMPLE: Using the Heat of Vaporization in Calculations

Calculate the amount of water, in grams, that can be vaporized at its boiling point with 155 kJ of heat.

GIVEN: 155 kJ FIND: g H_2O LOOK UP: ΔH_{vap}

SOLUTION MAP:

$$\text{kJ} \xrightarrow{\frac{1 \text{ mol } H_2O}{40.7 \text{ kJ}}} \text{mol } H_2O \xrightarrow{\frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O}} \text{g } H_2O$$

SOLUTION:

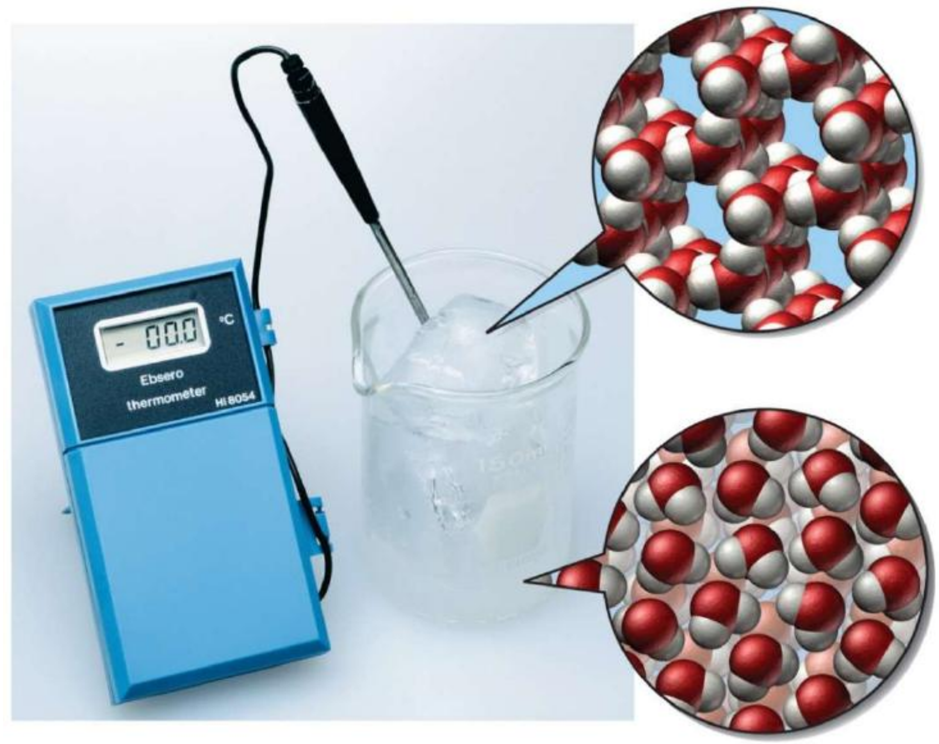
$$155\ \cancel{\text{kJ}} \times \frac{1\ \cancel{\text{mol } H_2O}}{40.7\ \cancel{\text{kJ}}} \times \frac{18.02 \text{ g}}{1\ \cancel{\text{mol } H_2O}} = 68.6 \text{ g}$$

Melting and Freezing

- As the temperature of a solid increases, thermal energy causes the molecules and atoms composing the solid to vibrate faster.
- At the **melting point,** atoms and molecules have enough thermal energy to overcome the intermolecular forces that hold them at their stationary points, and the solid turns into a liquid.

Melting and Freezing

- When ice melts, water molecules break free from the solid structure and become liquid.
- As long as ice and water are both present, the temperature will be 0 °C.

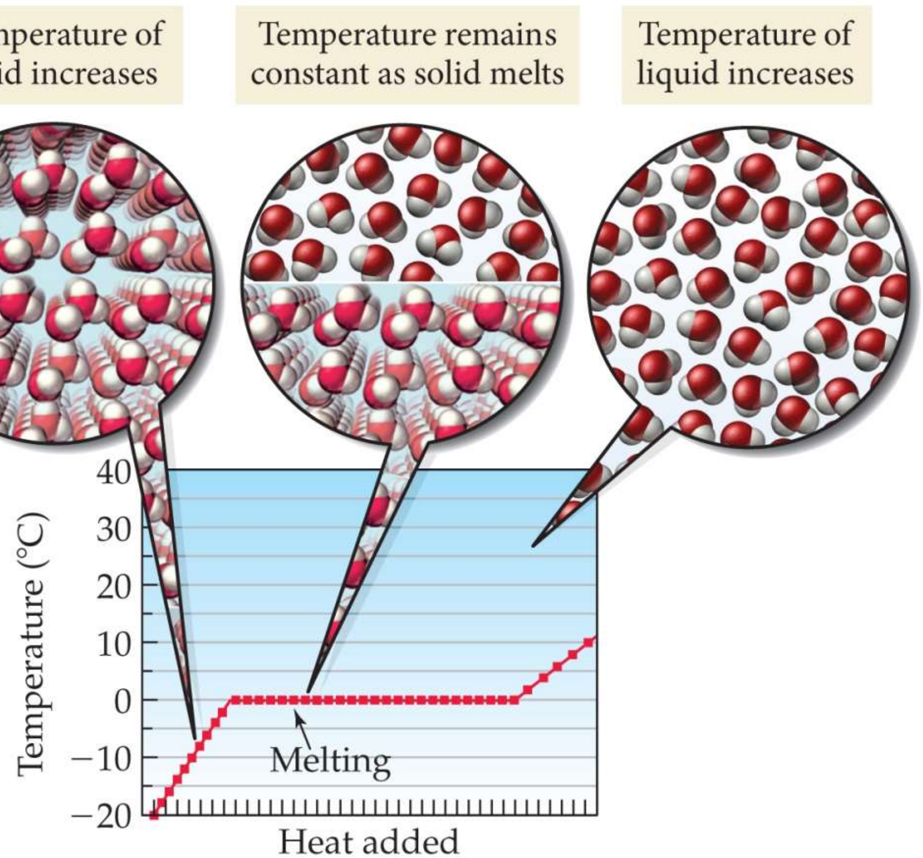


Remember: Phase transitions for a pure substance, at a given pressure, occur at a constant temperature

- A mixture of water *and* ice will always have a temperature of 0 °C (at 1 atm pressure).
- Only after all of the ice has melted will additional heating raise the temperature of the liquid water past 0 °C.

Another “Heating Curve” for Water

This time, a graph of the temperature of ice as it is heated from −20 °C to 35 °C (at the usual 1 atm)



During melting, the temperature of the solid and the liquid remains at 0 °C until the entire solid is melted.

Energetics of Melting and Freezing

- A way to cool down a drink is to drop ice cubes into it.
- As the ice melts, the drink cools because melting is endothermic—heat is absorbed when a solid is converted into a liquid. <u>Melting</u> of the ice absorbs heat.
- Melting is endothermic because energy is required to partially overcome the attractions between molecules in the solid and free them into the liquid state.
- Freezing is exothermic—heat is released when a liquid freezes into a solid.
- As water in a freezer turns into ice, it releases heat, which must be removed by the refrigeration system of the freezer.

Heat of Fusion

"Fusion": An old-fashioned term for melting

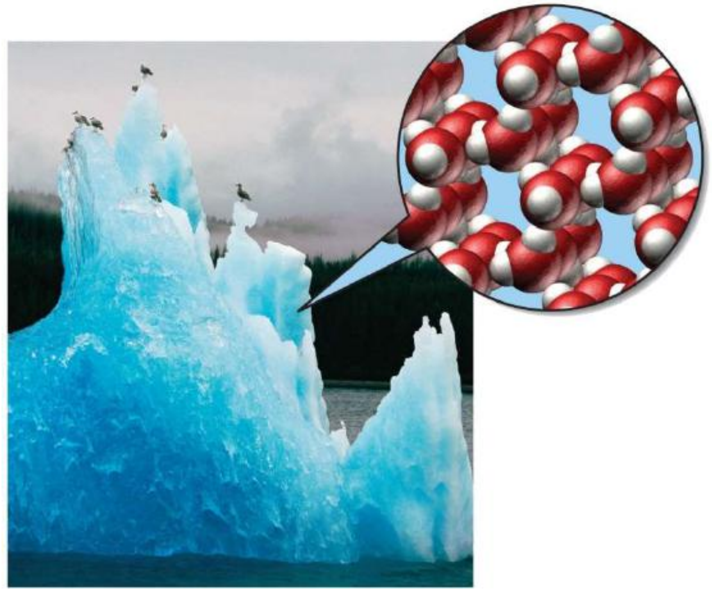
So, we are talking about "Heat of Melting" here

- The amount of heat required to melt 1 mol of a solid is the **heat of fusion (ΔH_{fus}).**
- The heat of fusion for water is 6.02 kJ/mol.
- ΔH_{fus} is positive because melting is endothermic; energy must be added to the ice to melt it.
- The same amount of heat is involved when 1 mol of liquid water freezes, but the *heat is emitted* rather than absorbed.
    - ΔH is negative because freezing is exothermic; energy is given off as the water freezes.

Heat of Fusion of Water

$$H_2O(s) \rightarrow H_2O(l) \qquad \Delta H = +6.02 \text{ kJ}$$

$$H_2O(l) \rightarrow H_2O(s) \qquad \Delta H = -6.02 \text{ kJ}$$



- Use the heat of fusion to calculate the amount of heat energy required to melt a given amount of a solid.
- Use the heat of fusion as a conversion factor between moles of a solid and the amount of heat required to melt them.

EXAMPLE: Using the Heat of Fusion in Calculations

Calculate the amount of ice in grams that, upon melting (at 0 °C), absorbs 237 kJ of heat.

GIVEN: 237 kJ   FIND: g H_2O LOOK UP: ΔH_{fus}

SOLUTION MAP:

$$\text{kJ} \longrightarrow \text{mol } H_2O \longrightarrow \text{g } H_2O$$

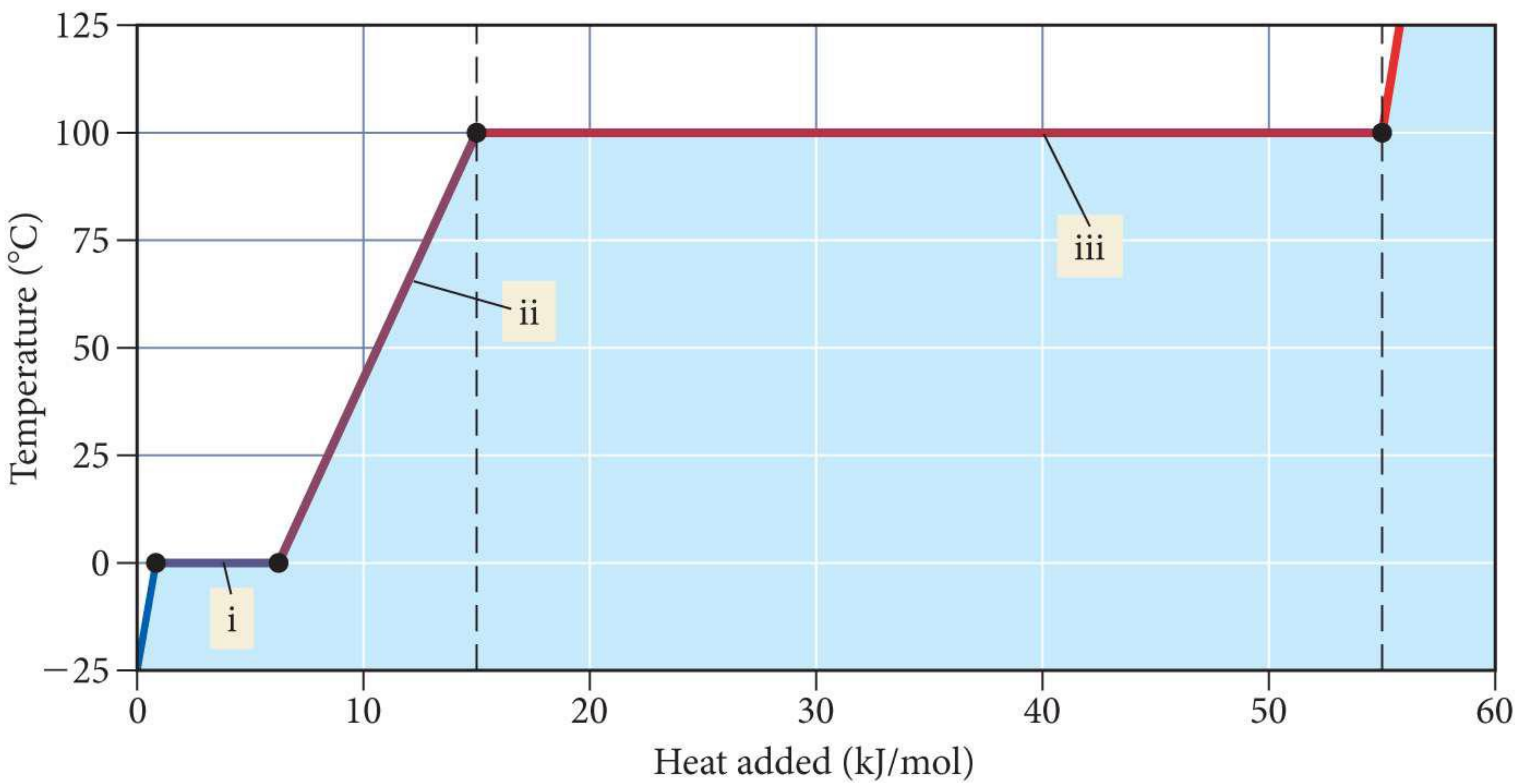
$$\frac{1 \text{ mol } H_2O}{6.02 \text{ kJ}} \qquad\qquad \frac{18.02 \text{ g}}{1 \text{ mol } H_2O}$$

SOLUTION:

$$237 \, \cancel{\text{kJ}} \times \frac{1 \, \cancel{\text{mol } H_2O}}{6.02 \, \cancel{\text{kJ}}} \times \frac{18.02 \text{ g}}{1 \, \cancel{\text{mol } H_2O}} = 709 \text{ g}$$

A Heating Curve for Ice

The diagram shows a heating curve for ice beginning at –25 °C and ending at 125 °C. Correlate sections i, ii, and iii with the correct states of water.

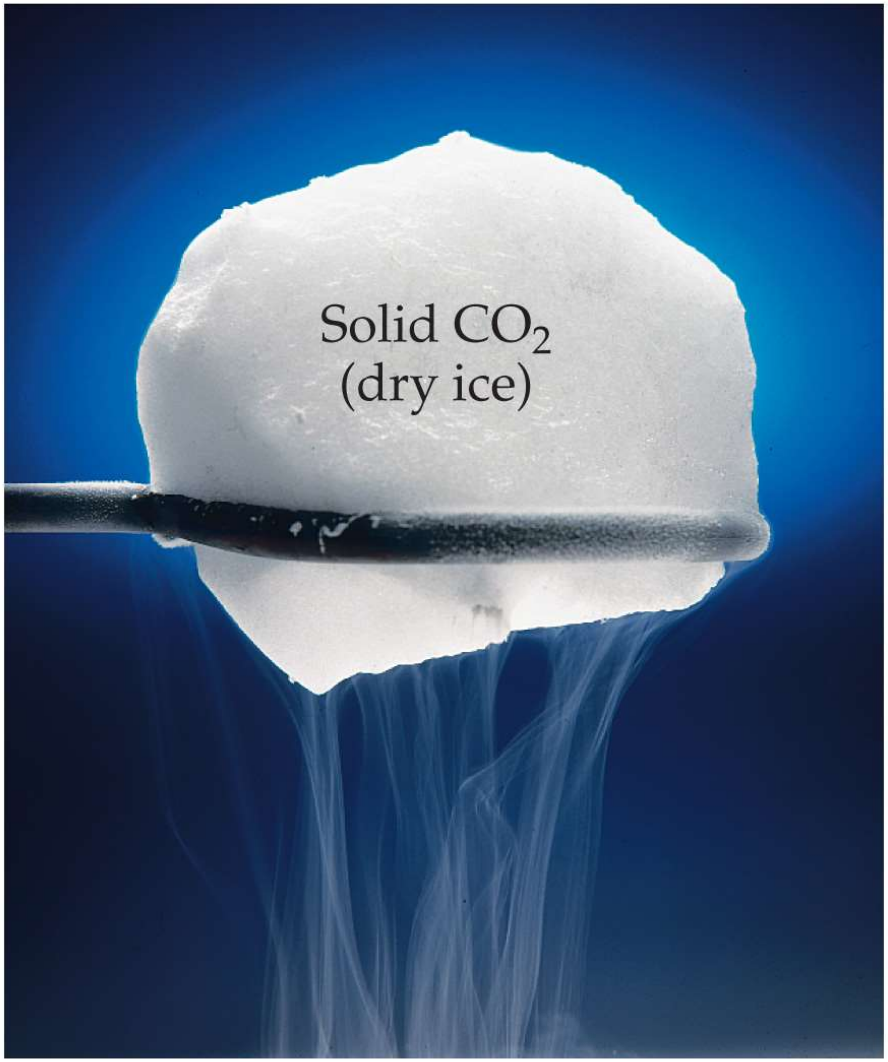


Sublimation

- **Sublimation** is a physical change in which a substance changes from its solid state directly to its gaseous state.
- When a substance sublimes, molecules leave the surface of the solid, where they are held less tightly than in the interior, and become gaseous.
- Dry ice, which is solid carbon dioxide, does not melt under atmospheric pressure (at any temperature).
- At −78 °C, the CO_2 molecules have enough energy to leave the surface of the dry ice and become gaseous, in a process analogous to boiling
  - Its vapor pressure is equal to 1 atm
- Solids sublime to some extent at every temperature, just like liquids. It's usually a lot less, so we just don't notice it.

Sublimation

- Dry ice is solid carbon dioxide. The solid does not melt but rather sublimes. It transforms directly from solid carbon dioxide to gaseous carbon dioxide.



Sublimation

- Regular ice will slowly sublime at temperatures below 0 °C.
- In cold climates, ice or snow lying on the ground gradually disappears, even if the temperature remains below 0 °C.
- Similarly, ice cubes left in the freezer for a long time slowly become smaller, even though the freezer is always below 0 °C.

Sublimation

- Ice sublimes out of frozen foods.
- You can see this in food that has been frozen in an airtight plastic bag for a long time.
- The ice crystals that form in the bag are water that has sublimed out of the food and redeposited on the surface of the bag.
- Food that remains frozen for too long becomes dried out.
- This can be avoided by freezing foods to colder temperatures (further below 0 °C), a process called deep-freezing.
- The colder temperature lowers the rate of sublimation and preserves the food longer.

Intermolecular forces also affect whether two substances will be soluble in one another

If there were no intermolecular forces, everything would freely dissolve in everything else.

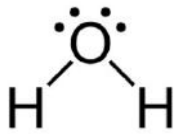
If dissolution of two substances would disrupt the intermolecular forces, their energy would go up.

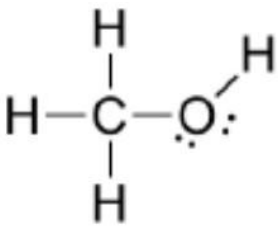
- Dissolving unfavorable; little solubility

If the two substances have intermolecular forces similar in nature, they tend to be "miscible"

- they dissolve in each other in a wide range of ratios

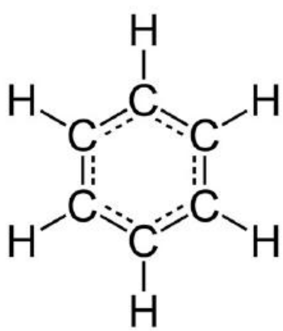
"like dissolves like"



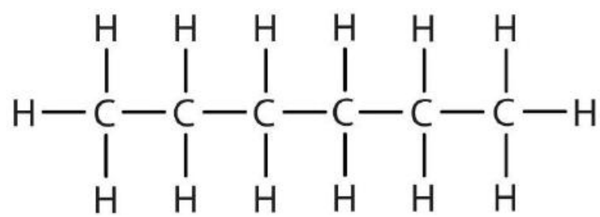


H_2O and CH_3OH both have dipole-dipole and hydrogen bonding interactions

– Miscible (can dissolve in all proportions)

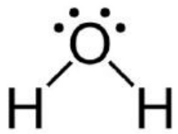


and

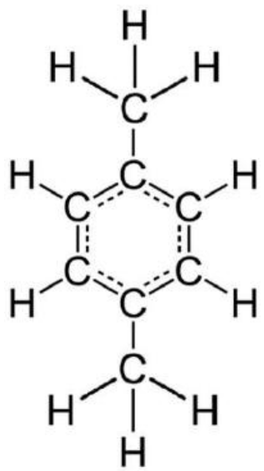


both have basically only London dispersion forces

– Miscible (can dissolve in all proportions)



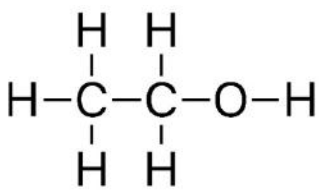
H_2O has strong hydrogen bonding forces and significant dipole-dipole forces, but very little dispersion



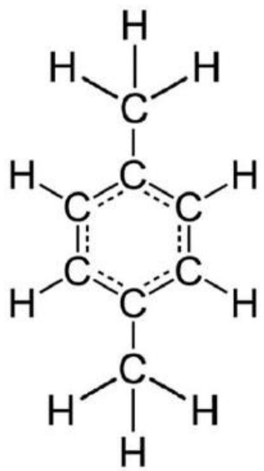
para-Xylene has strong London dispersion forces (actually a bit stronger than the forces between H_2O molecules), but no other forces

Water and para-Xylene are **insoluble in each other**

Dissolution is especially bad for water:
H-bonding and dipole-dipole interactions would be lost, and it simply doesn’t have enough electrons to have strong London-dispersion forces to replace what’s lost.



Ethanol has hydrogen bonding, dipole-dipole (it is asymmetric, therefore polar) and significant dispersion forces.



para-Xylene has strong London dispersion forces, and no other forces

Ethanol and para-Xylene are **miscible**

Since ethanol has a strong component of dispersion forces between its molecules, it is compatible enough with a substance that has only dispersion forces.

Ethanol can be seen as a more versatile solvent than water.

Practice

Choose the substance in that boils at a higher temperature

Check your answer on the next page

a) CH_3OH CH_3CHF_2

b) CH_3-O-CH_2CH_3 $CH_3CH_2CH_2NH_2$

Practice

Choose the substance in that boils at a higher temperature

a) CH_3OH CH_3CHF_2

H-bonding
Dipole-dipole

b) $CH_3\text{-}O\text{-}CH_2CH_3$ $CH_3CH_2CH_2NH_2$

can H-bond

Practice

Choose the substance in that boils at a higher temperature

Check your answer on the next page

CH_3OH CH_3F

Practice

Choose the substance in that boils at a higher temperature

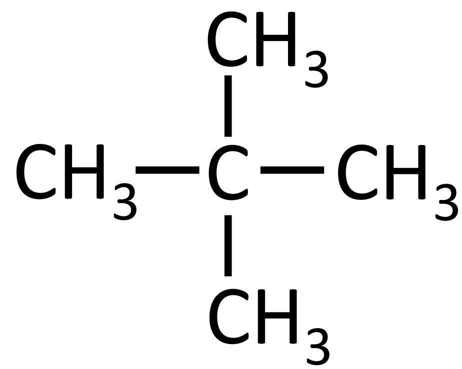
CH_3OH		CH_3F
H-bonding		No H-bonding
Dipole-dipole	⟷ comparable	Dipole-dipole
Dispersion (by 18 e^-)	⟷ comparable	Dispersion (by 18 e^-)

Practice

Choose the substance in that boils at a higher temperature

Check your answer on the next page

CH_3CH_3



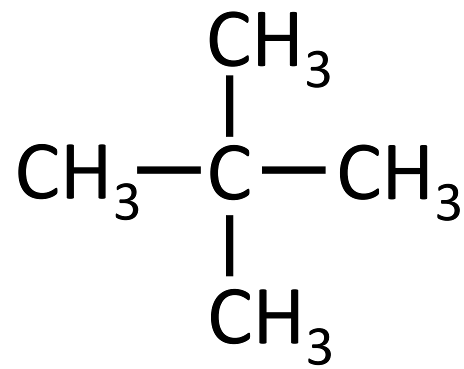
Practice

Choose the substance in that boils at a higher temperature

CH_3CH_3

No dipole-dipole
(symmetric)

Dispersion (by 18 e^-)



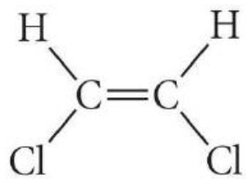
No dipole-dipole
(symmetric)

Dispersion (by 72 e^-)

Practice

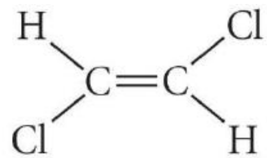
Choose the substance in that boils at a higher temperature

Check your answer on the next page



cis-1,2-Dichloroethene

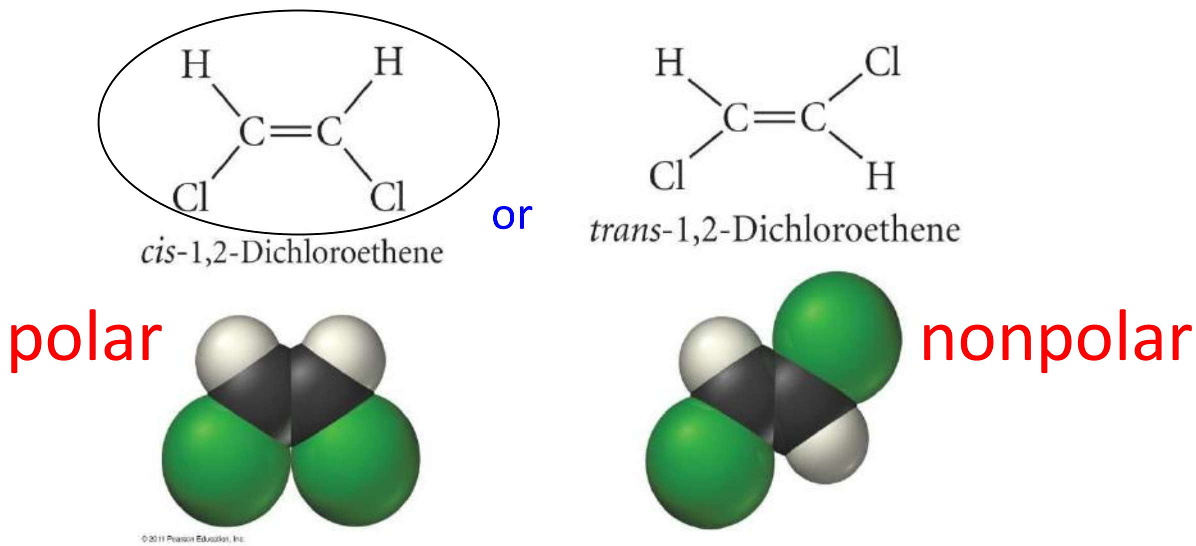
or



trans-1,2-Dichloroethene

Practice

Choose the substance in that boils at a higher temperature



Both lack hydrogen-bonding

Both have the same number of electrons & comparable size

The polar molecule has dipole-dipole forces, whereas the nonpolar molecule doesn't

Chemistry and Life:
Hydrogen Bonding in DNA

- A DNA molecule is composed of thousands of repeating units called *nucleotides*.
- Each nucleotide contains a base: adenine, thymine, cytosine, or guanine (abbreviated A, T, C, and G).
- The order of these bases along DNA encodes the instructions that specify how proteins are made in each cell of the body.
- DNA consists of two complementary strands wrapped around each other in the now-famous double helix.
- Each strand is held to the other by ***hydrogen bonds*** that occur between the bases on each strand.

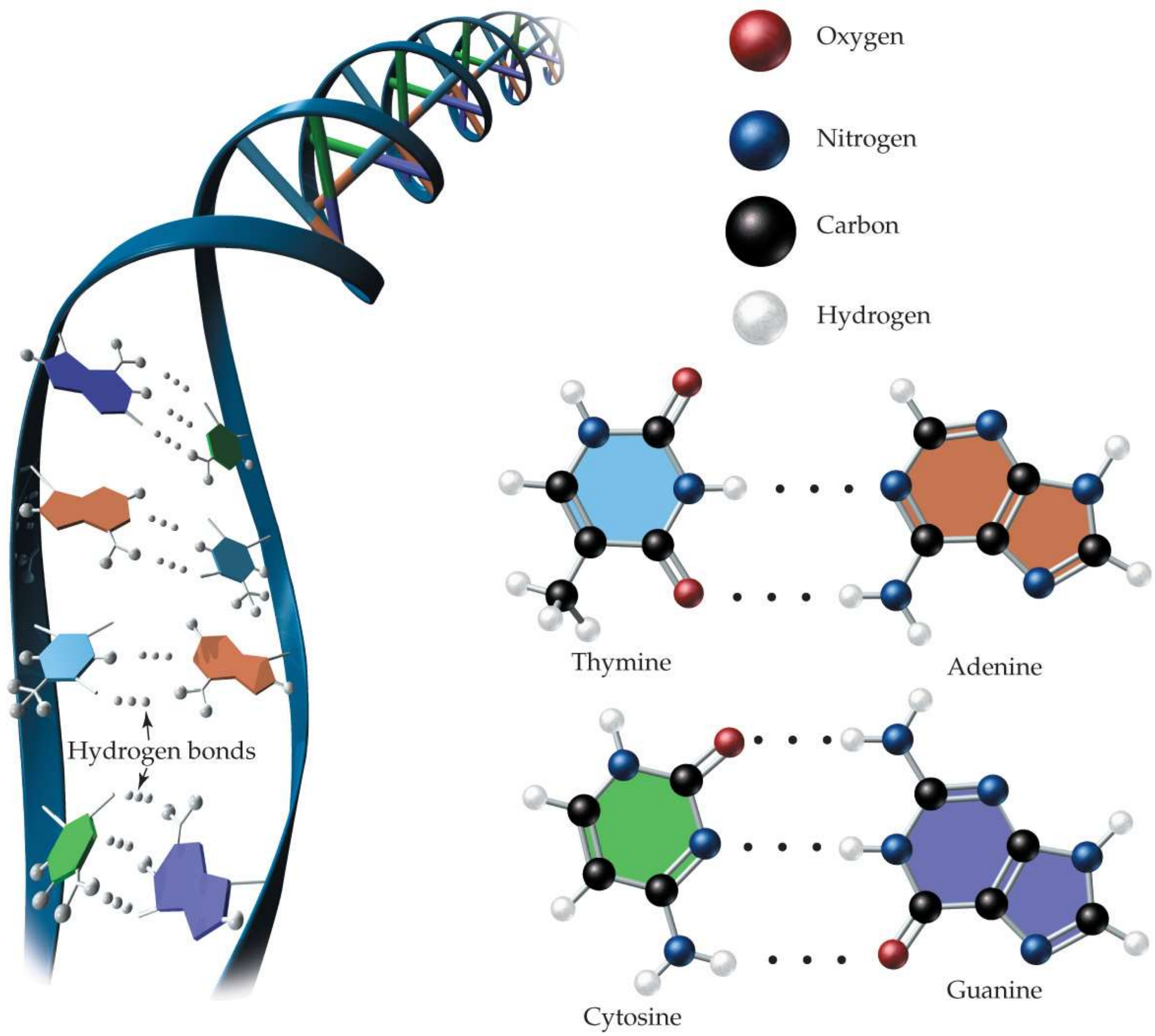
Chemistry and Life: Hydrogen Bonding in DNA

- DNA replicates because each base (A, T, C, and G) has a complementary partner with which it hydrogen-bonds.
- Adenine (A) hydrogen-bonds with thymine (T).
- Cytosine (C) hydrogen-bonds with guanine (G).
- The ***hydrogen bonds*** are so specific that each base will pair only with its complementary partner.

Chemistry and Health: Hydrogen Bonding in DNA

- When a cell is going to divide, the DNA unzips across the hydrogen bonds that run along its length.
- New bases, complementary to the bases in each half, add along each of the halves, forming ***hydrogen bonds*** with their complement.
- The result is two identical copies of the original DNA.

Hydrogen Bonding in DNA



Types of Crystalline Solids: Molecular, Ionic, and Atomic

- Solids may be crystalline (a well-ordered array of atoms or molecules) or amorphous (having no long-range order).
- Crystalline solids can be divided into three categories—molecular, ionic, and atomic—based on the individual units that compose the solid.

Types of Crystalline Solids

Crystalline solids

Molecular solids

Composite units are molecules

Low melting points

Ice

but can be > 300 °C for large molecules

Ionic solids

Composite units are formula units (cations and anions)

High melting points

(not always)

Table salt

but also: "ionic liquids" (liquid at room temperature) with large, odd-shaped ions

Atomic solids

Composite units are atoms

Variable melting points

(very low to very high)

Gold

but also: Helium (doesn't freeze even at 0 Kelvin at 1 atm

and also: Carbon (more than 4000 K, at high pressure)

Molecular Solids

- **Molecular solids** are solids whose composite units are *molecules*.
- Ice (solid H_2O) and dry ice (solid CO_2) are examples of molecular solids.
- Molecular solids are held together by intermolecular forces—dispersion forces, dipole–dipole forces, and hydrogen bonding.
- Ice is held together by hydrogen bonds, and dry ice is held together by dispersion forces.
- Molecular solids as a whole tend to have low to moderately low melting points.
- Ice melts at 0 °C and dry ice sublimes at −78 °C.

Ionic Solids

- **Ionic solids** are solids composed of *formula units,* the smallest electrically neutral collection of cations and anions that compose the compound.
- Table salt (NaCl) and calcium fluoride (CaF_2) are good examples of ionic solids.
- Ionic solids are held together by electrostatic attractions between cations and anions.
- In NaCl, the attraction between the Na^+ cation and the Cl^- anion holds the solid lattice together because the lattice is composed of alternating cations and anions in a three-dimensional array.

Ionic Solids

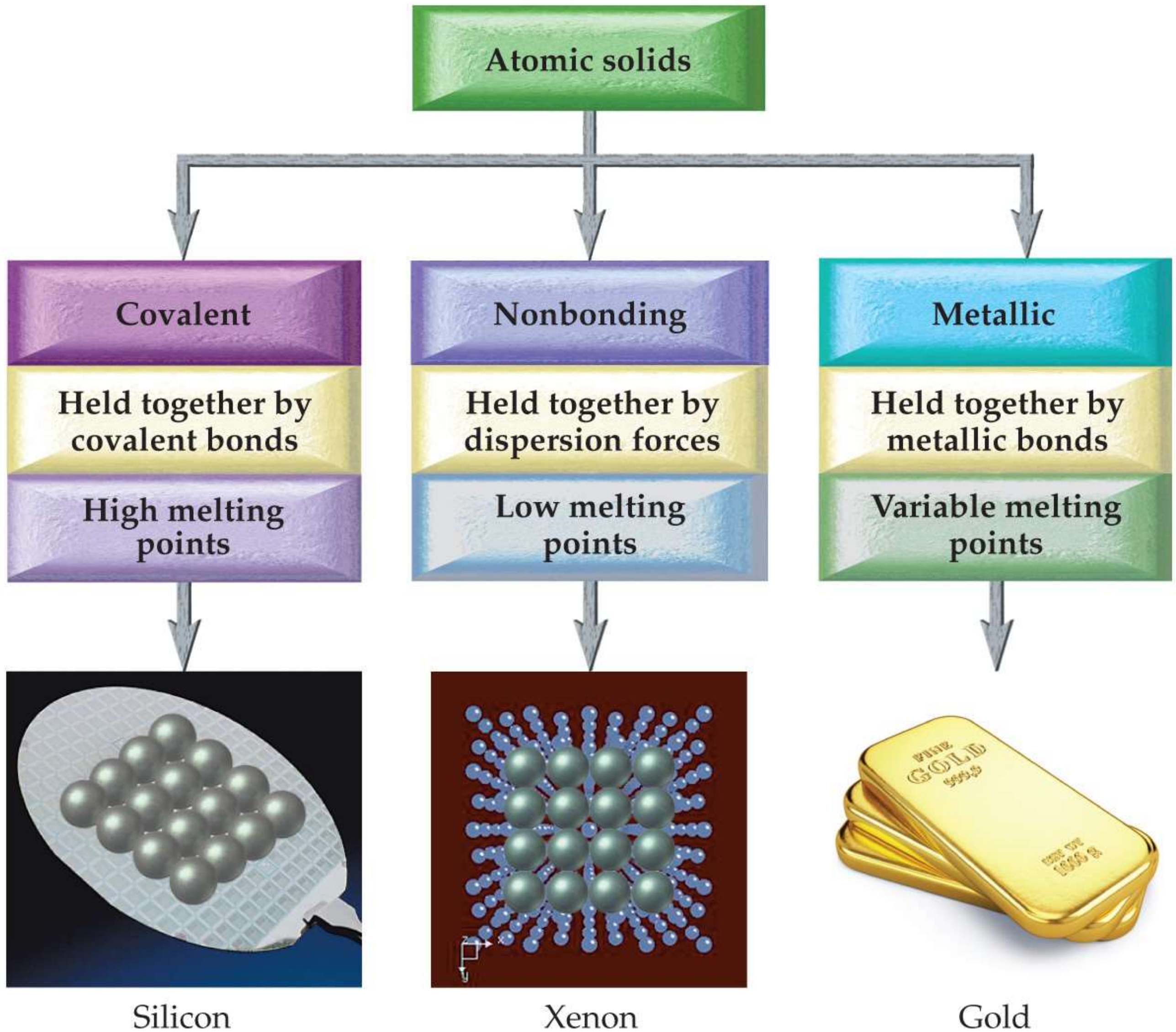
- The forces that hold ionic solids together are actual ionic bonds.
- Since *ionic bonds are* much *stronger than* any of the *intermolecular forces* discussed previously, ionic solids tend to have much higher melting points than molecular solids. *
- Sodium chloride, an ionic solid, melts at 801 °C, while carbon disulfide, a molecular solid with a higher molar mass, melts at −110 °C. *

* These observations mainly apply to small molecules and simple ionic compounds
There are many molecular compounds that melt higher than many “ionic liquids”

Atomic Solids

- **Atomic solids** are solids whose composite units are *individual atoms*.
- Diamond (C), iron (Fe), and solid xenon (Xe) are good examples of atomic solids.
- Atomic solids can be divided into three categories—**covalent atomic solids, nonbonding atomic solids,** and **metallic atomic solids**—each held together by a different kind of force.

Types of Atomic Solids

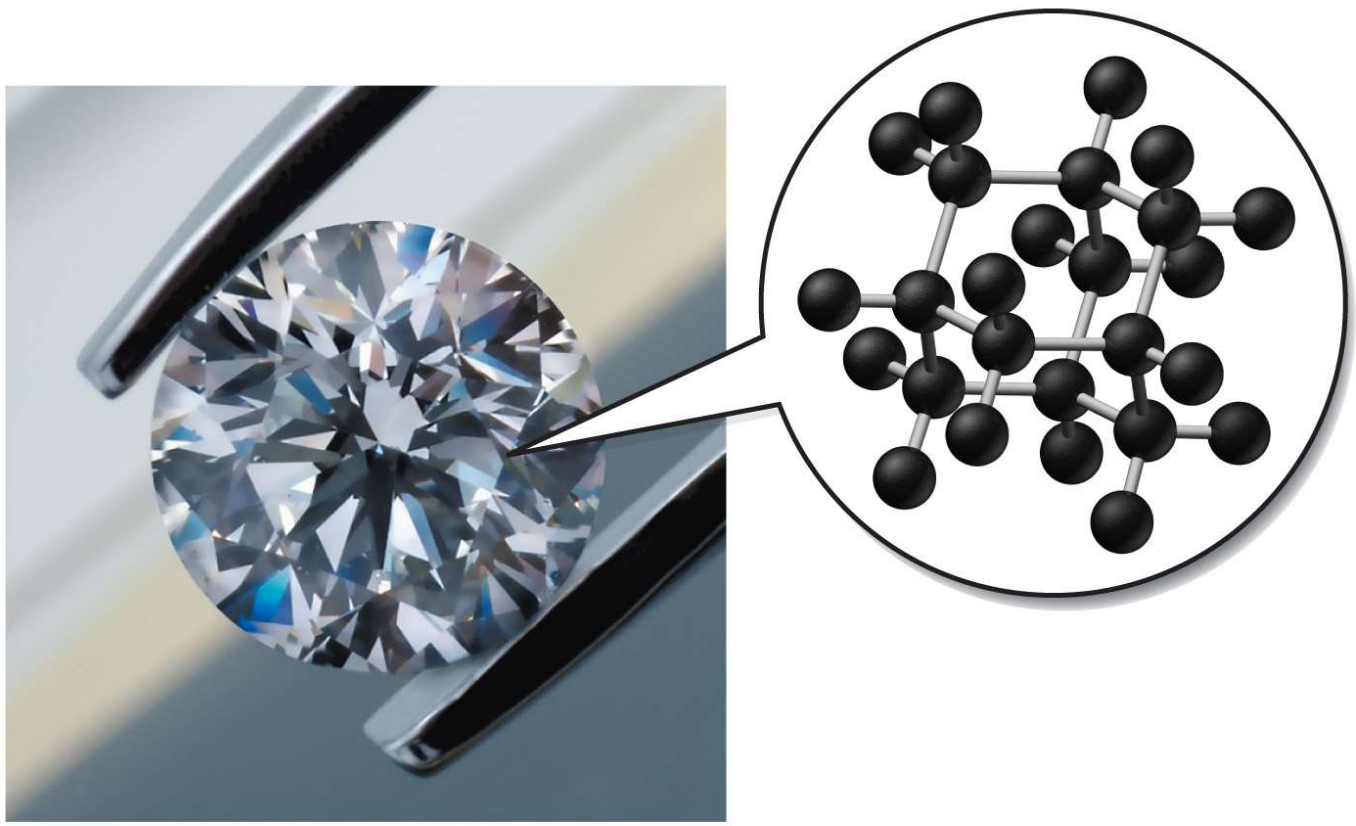


Covalent Atomic Solids

- Covalent atomic solids, such as diamond, are held together by covalent bonds.
- In diamond, each carbon atom forms four covalent bonds to four other carbon atoms in a tetrahedral geometry.
- This structure extends throughout the entire crystal, so that a diamond crystal can be thought of as a giant molecule held together by these covalent bonds.
- Since covalent bonds are very strong, covalent atomic solids have high melting points. Diamond is estimated to melt at about 3800 °C.

Diamond: A Covalent Atomic Solid

In diamond, carbon atoms form covalent bonds in a three-dimensional hexagonal pattern.



Nonbonding Atomic Solids

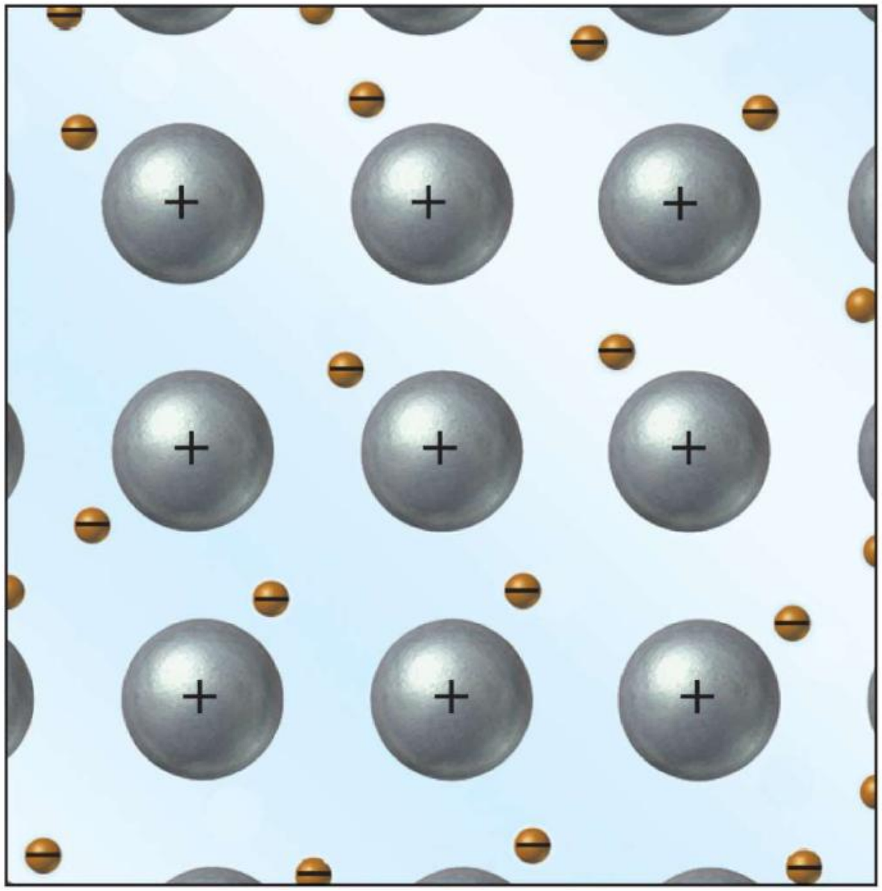
- Nonbonding atomic solids, such as solid xenon, are held together by relatively weak dispersion forces.
- Xenon atoms have stable electron configurations and therefore do not form covalent bonds with each other.
- Consequently, solid xenon, like other nonbonding atomic solids, has a very low melting point (about $-112\ ^{\circ}C$).

Metallic Atomic Solids

- Metallic atomic solids, such as iron, silver, and lead, have variable melting points.
- Metals are held together by metallic bonds that, in the simplest model, consist of positively charged ions in a sea of electrons.
- Metallic bonds are of varying strengths, with some metals, such as mercury, having melting points below room temperature (−39 °C) and other metals, such as iron, having relatively high melting points (iron melts at 1809 °C).

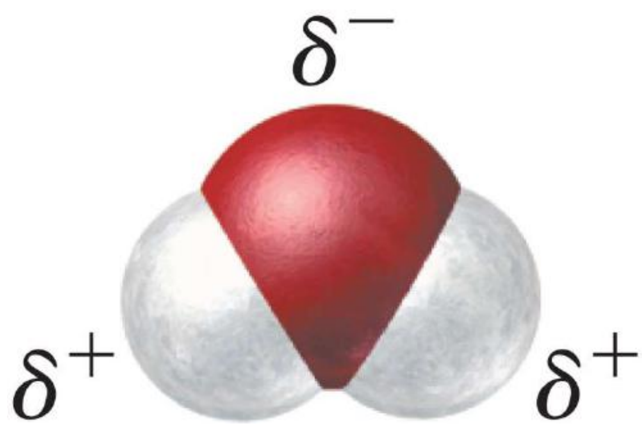
Structure of a Metallic Atomic Solid

- In the simplest model of a metal, each atom donates one or more electrons to an "electron sea."
- The metal consists of the metal cations in a negatively charged electron sea.



Water: A Remarkable Molecule

- Water has a low molar mass (18.02 g/mol), yet it is a liquid instead of a gas at room temperature.
- Water's relatively high boiling point can be understood by examining the structure of the water molecule.



Water: A Remarkable Molecule

- The bent geometry of the water molecule and the highly polar nature of the O — H bonds result in a molecule with a significant dipole moment.
- Water’s two O — H bonds (hydrogen directly bonded to oxygen) allow water molecules to form strong hydrogen bonds with other water molecules, resulting in a relatively high boiling point.
- Water’s high polarity also allows it to dissolve many other polar and ionic compounds.
- Consequently, water is the main solvent of living organisms, transporting nutrients and other important compounds throughout the body.

Water: A Remarkable Molecule

- Life is impossible without water, and in most places on Earth where liquid water exists, life exists.
- Recent evidence of water on Mars—that either existed in the past or exists in the present—has fueled hopes of finding life or evidence of life there.

Water: A Remarkable Molecule

- The way water freezes is unique. Most other substances contract upon freezing, but water expands upon freezing.
- Because liquid water expands when it freezes, ice is less dense than liquid water. Water reaches its maximum density at 4.0 °C.
- Consequently, ice cubes and icebergs float.
- The frozen layer of ice at the surface of a winter lake insulates the water in the lake from further freezing.
- If ice were to sink, liquid water would remain exposed and keep freezing. Life in the body of water would die.
- Life survived extreme ice ages (“snowball earth”) because of this! Otherwise even the oceans would have frozen solid!

Water: A Remarkable Molecule

- The expansion of water upon freezing is one reason that most organisms do not survive freezing.
- When the water within a cell freezes, it expands and often ruptures the cell, just as water freezing within a pipe bursts the pipe.
- Many foods, especially those with high water content, do not survive freezing very well either.
- Industrial flash-freezing of fruits and vegetables happens so rapidly that the water molecules cannot align into the expanded phase and so the cells are not ruptured.