Liquids, Solids,
&
Intermolecular Forces

Intermolecular forces

- London dispersion forces
- Dipole-dipole interaction
- Ion-dipole interaction
- Hydrogen bonding

$<$ **ionic bond**
covalent bond

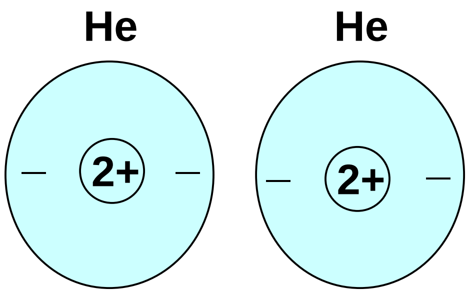
↓

Intramolecular forces

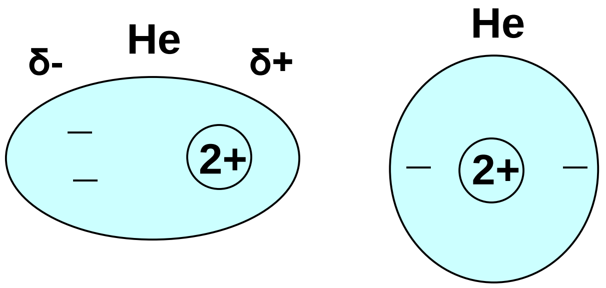
London dispersion forces

Attractive forces between all molecules

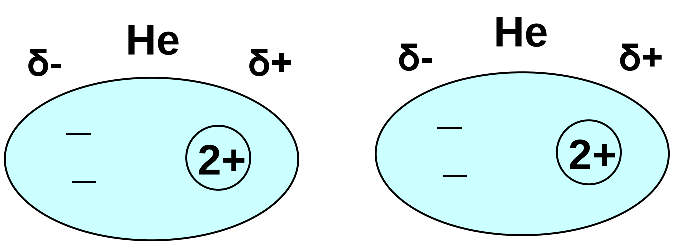
Only forces between nonpolar covalent molecules



No Polarity



Original Temporary Dipole



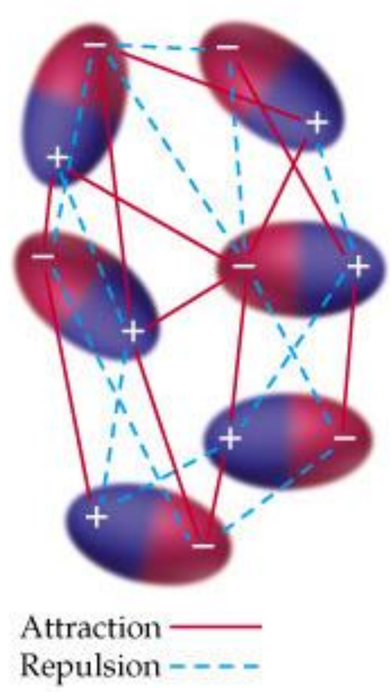
Original Temporary Dipole

Induced Temporary Dipole

London dispersion forces

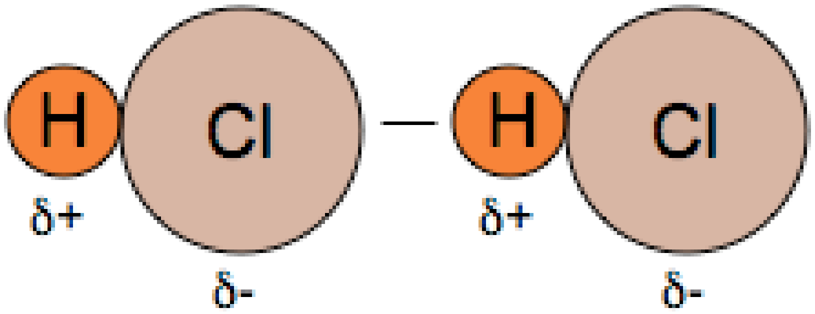
He: T = -240°C (1 atm) → liquid

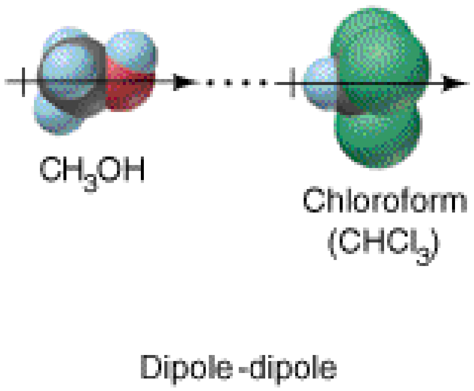
T ↓ ⇨ Kinetic energy ↓ Move slower ⇨ Attractive forces become more important ⇨ liquid



Dipole-Dipole Interactions

Attractive force between two polar molecules.





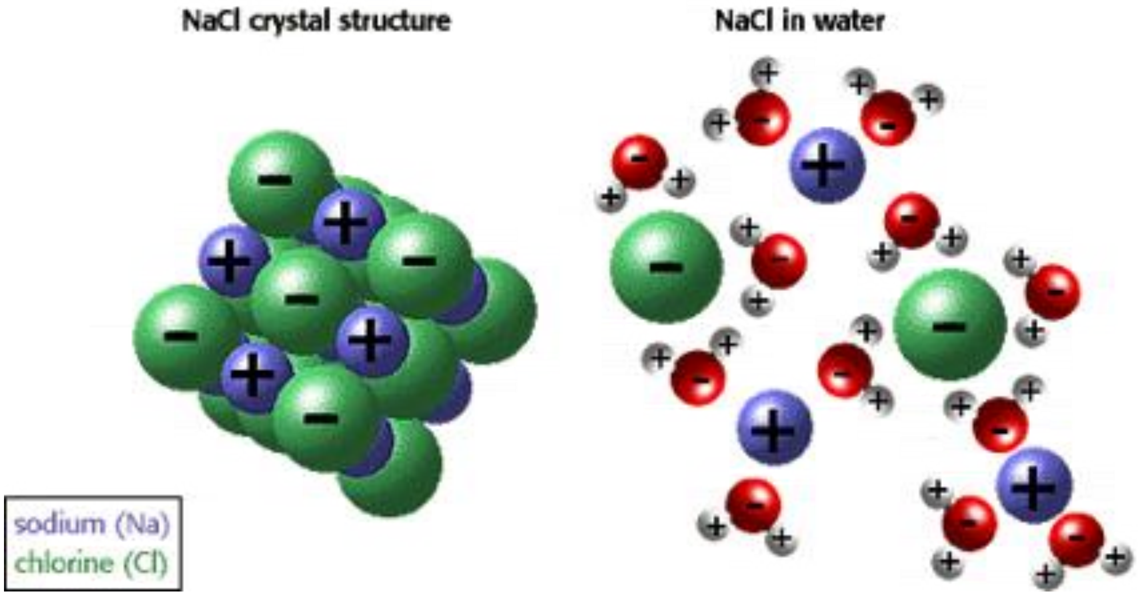
stronger than London dispersion forces

↓

↑ intermolecular forces → ↑ boiling point

Ion-Dipole Interactions

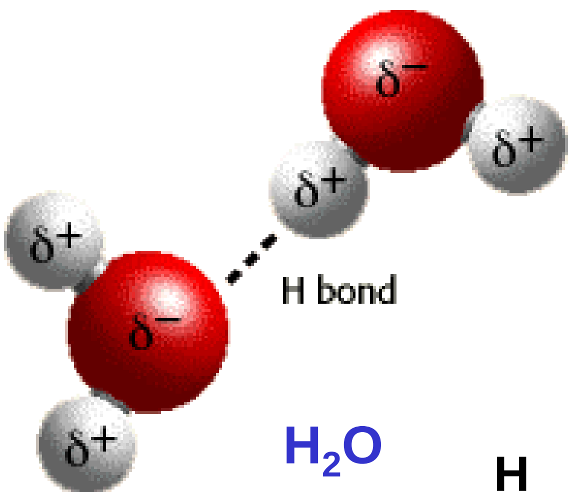
Attractive force between ionic compounds and polar molecules.



Very strong attraction.

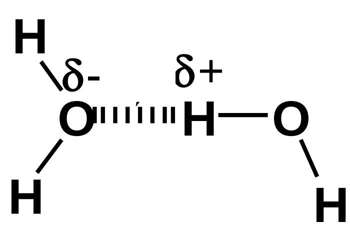
Hydrogen Bonds

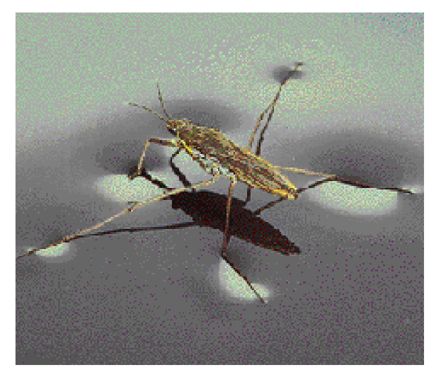
Between H bonded to O, N, or F (high electronegativity) → δ+
and a nearby O, N, or F → δ-



H_2O

Stronger than dipole-dipole interactions
& London dispersion forces





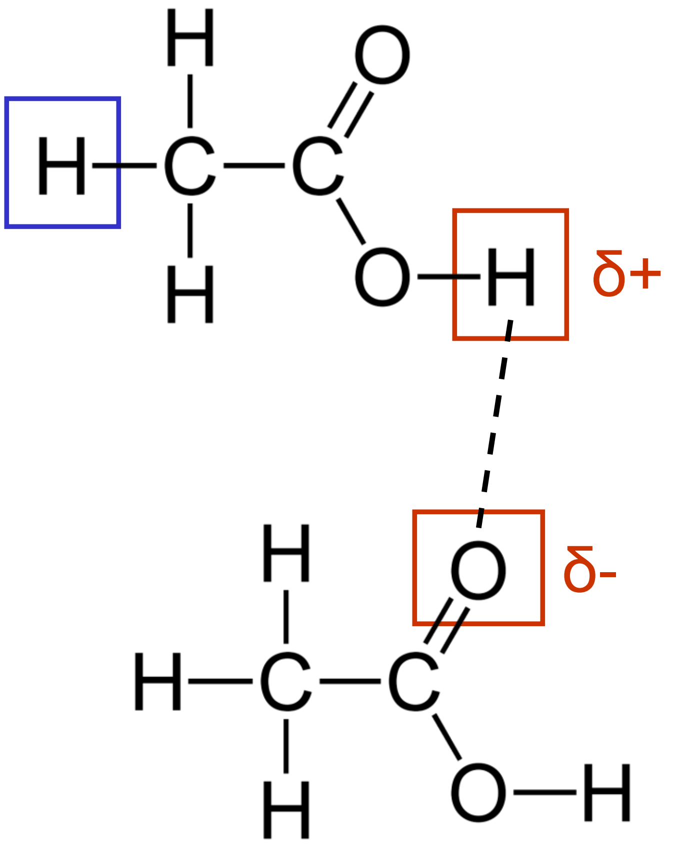
High boiling point

surface tension

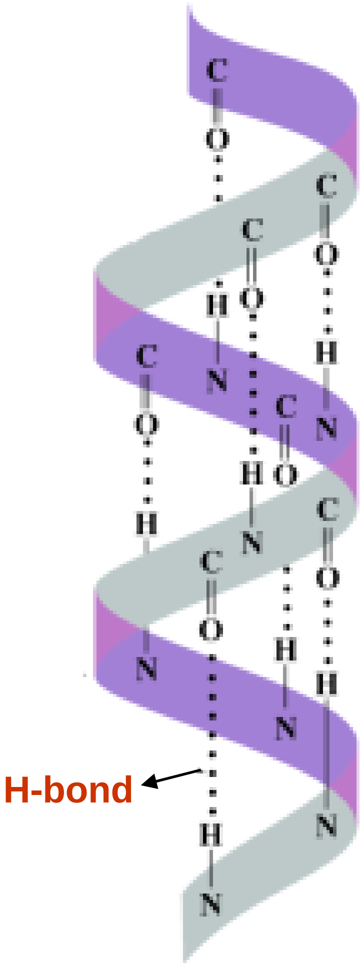
Hydrogen bonding

CH_3COOH

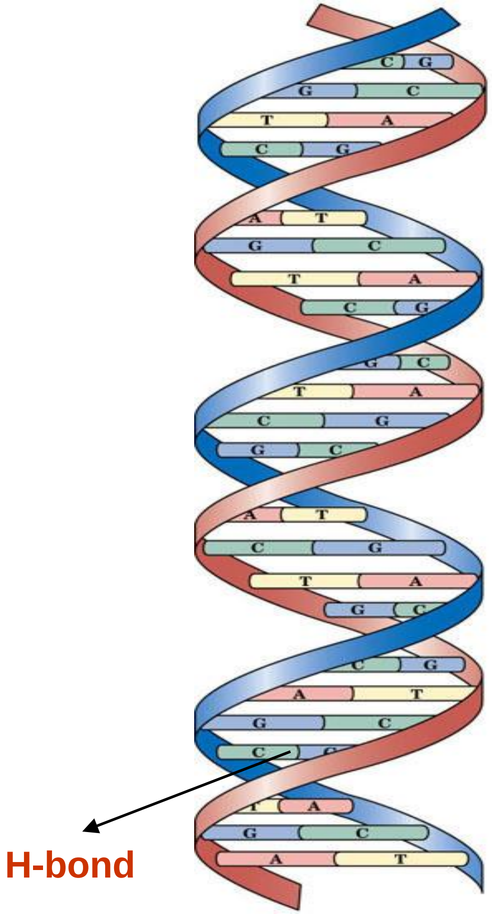
Acetic acid

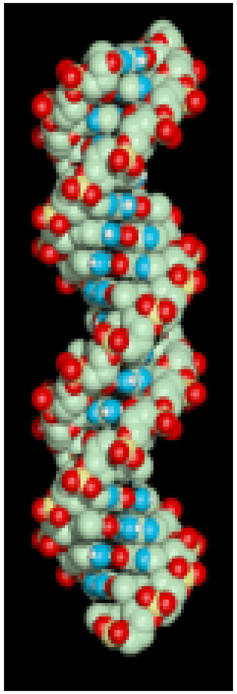


H-bonding in our body



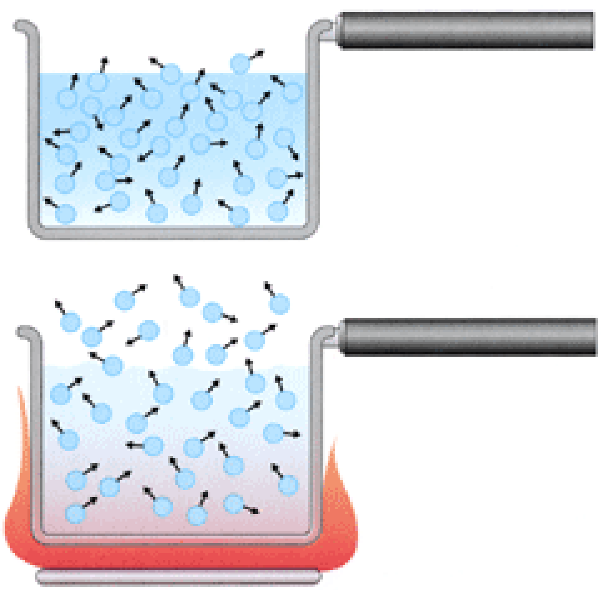
Protein (α-helix)





DNA

Evaporation

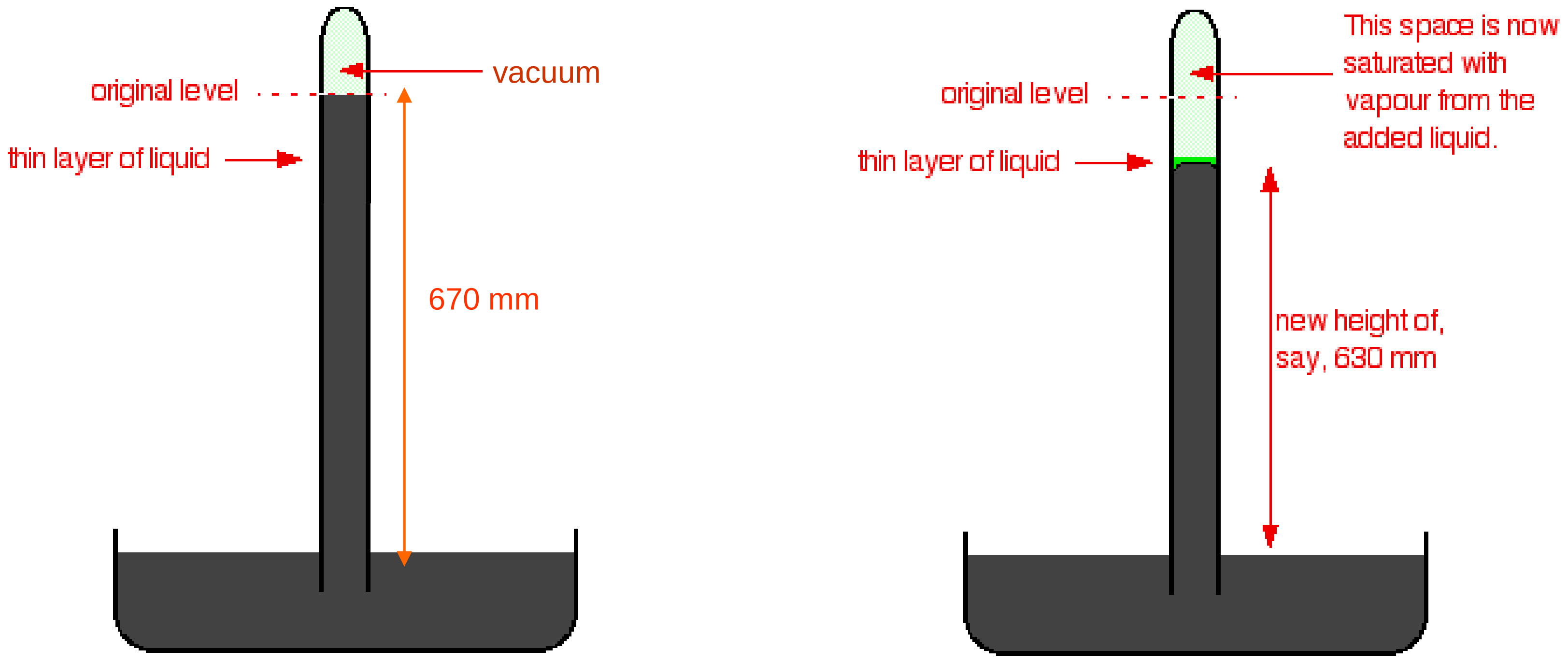


equilibrium

Vapor pressure: the pressure of a gas in equilibrium with its liquid form in a closed container.

Boiling point: the temperature at which the vapor pressure of a liquid is equal to the atmospheric pressure.

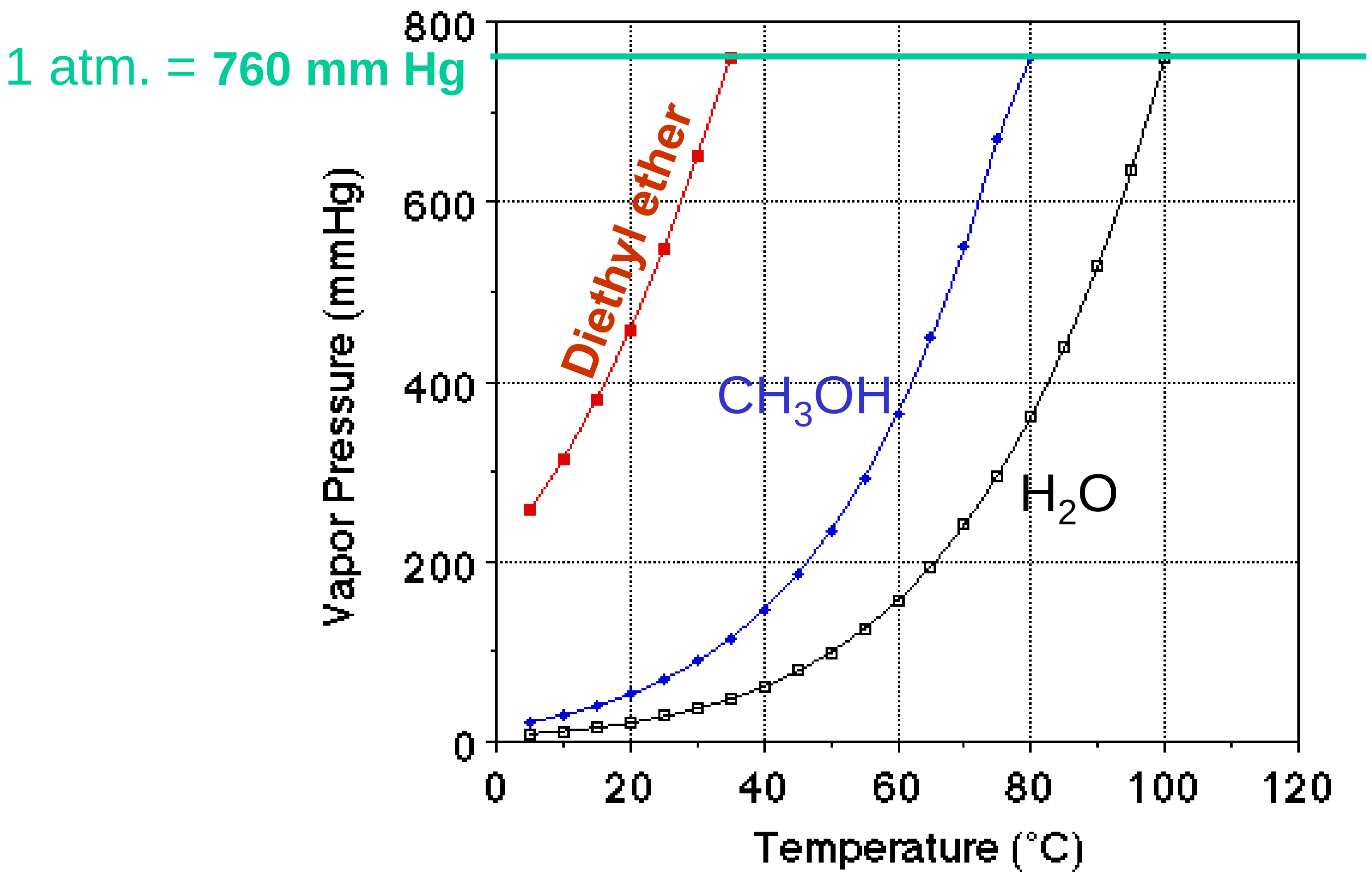
Evaporation



Measuring vapor pressure of liquids

Evaporation

normal boiling point: the temperature at which a liquid boils under a pressure of 1.00 atm.



Evaporation

Factors that affect boiling point:

1. Intermolecular forces:

London dispersion forces < Dipole-Dipole Int. < H-bonding < Ion-Dipole Int.

2. Number of sites for intermolecular interactions (surface area):

Larger surface (more electrons) $\rightarrow$ more sites for London $\rightarrow \uparrow$ b.p.

$CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_2\text{-}CH_3 > CH_3\text{-}CH_2\text{-}CH_3$

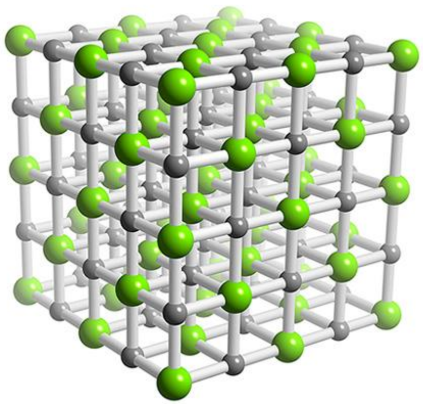
3. Molecular shape: With the same molecular weight.

linear $CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_2\text{-}CH_3$ > spherical

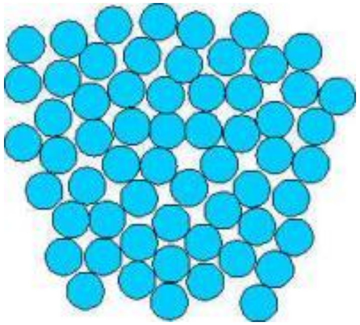
$$\begin{array}{c} CH_3 \\ | \\ CH_3\text{-}C\text{-}CH_3 \\ | \\ CH_3 \end{array}$$

Solid

Crystalline solid (Network solids)



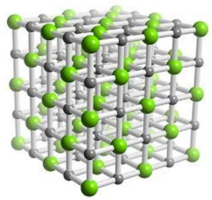
Amorphous solid



Crystalline solids (Network solids)

Crystalline solid

- Ionic solids: Consist of ions (metal-nonmetal) NaCl



Stable - High melting points

- Molecular solids: Consist of molecules. Sugar, Ice

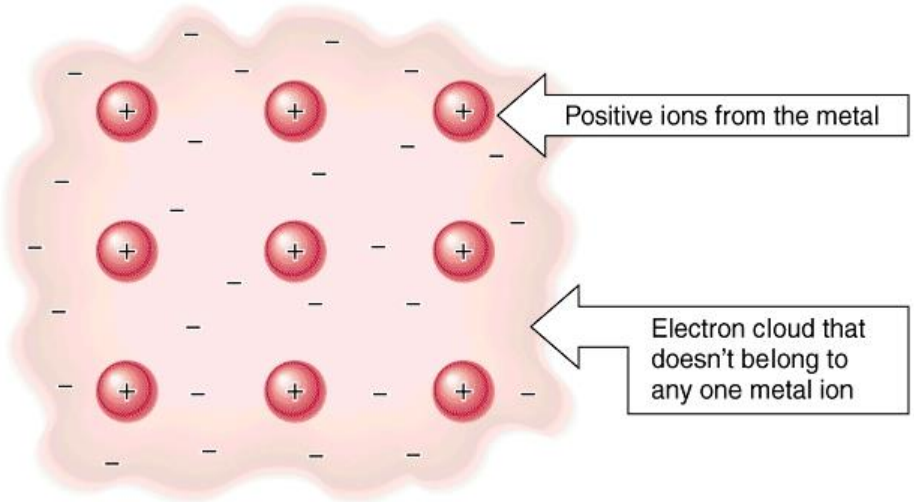
Lower melting points

London dispersion forces, Dipole-Dipole interaction, H-Bond

- Atomic solids: Consist of atoms. Diamond, Graphite, Metals

Different melting points (because of forces between atoms).

Bonding in metals



Electron Sea Model

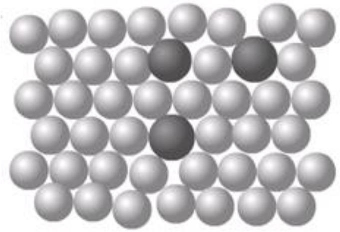
Valance electrons are shared among the atoms in a nondirectional way.

Metals conduct heat and electricity.

They are malleable and ductile.

We can make alloys.

Alloys



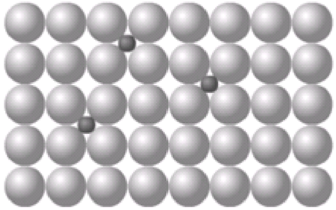
Substitutional alloy:

Some of the host metal atoms are replaced by other metal atoms of similar sizes.

Brass: (Copper, Zinc)

Interstitial alloy: Some of the holes among the metal atoms are occupied by atoms much smaller.

Steel: (Iron, Carbon)



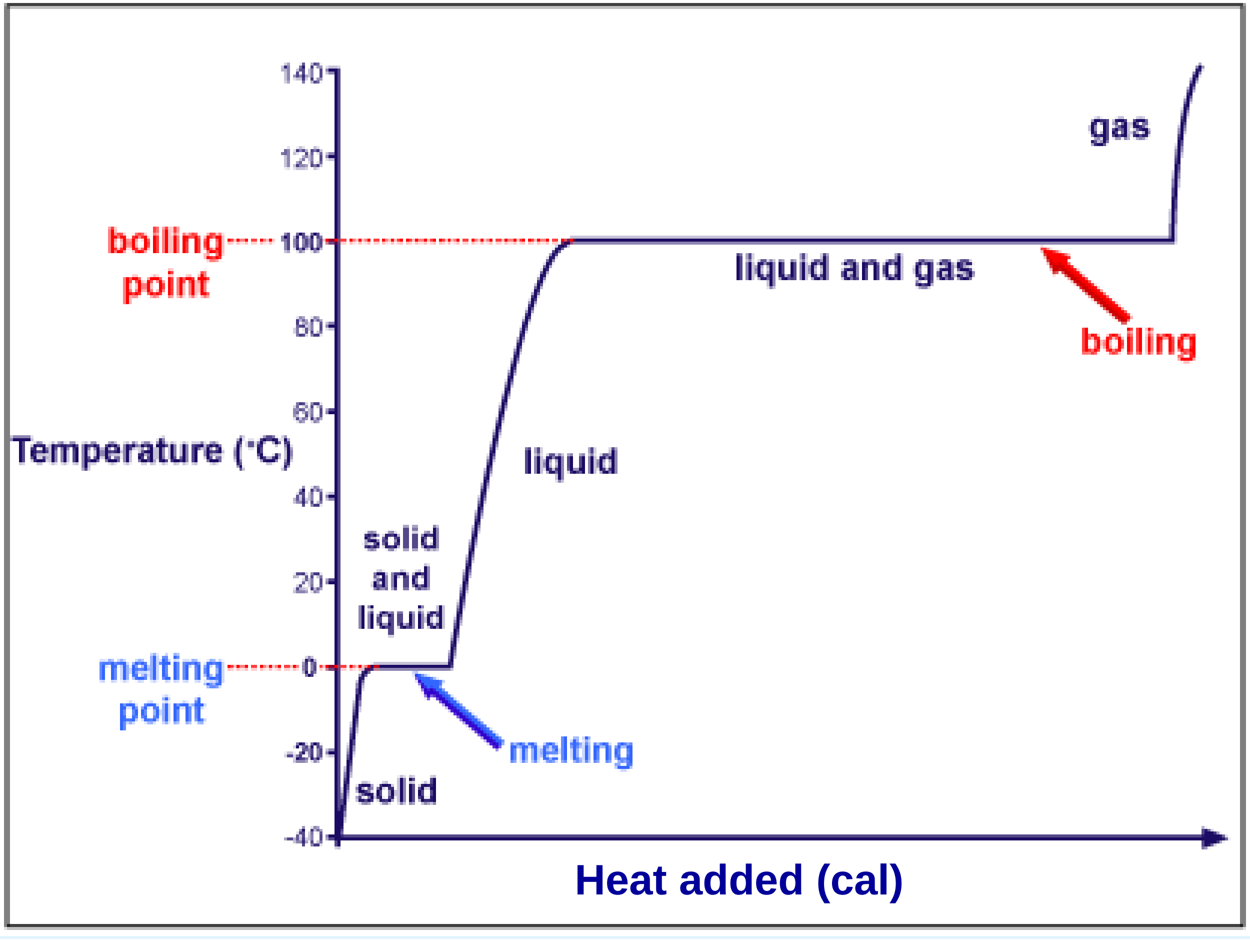
Solid

Solidification (Crystallization): change phase from liquid to solid.

Fusion (Melting): change phase from solid to liquid.

Sublimation: change phase from solid directly into the vapor.

Heating/Cooling Curve



during the phase changes, the temperature stays constant.

Heat and physical state

Molar heat of fusion: Energy required to melt 1 mol of a solid.

(For ice: 6.02 kJ/mol)

Molar heat of vaporization: Energy required to vaporize 1 mol of liquid.

(For water: 40.6 kJ/mol)

We need more energy for vaporization than fusion: Why?

To separate molecules enough to form a gas all of the intermolecular forces must be overcome.

Example 1

Ex. 12.2 Page 422:

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

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

Example 2

Calculate the amount of heat required to melt 25.0 g of ice (at 0ºC).