Intermolecular forces and Solutions
Chapter 7

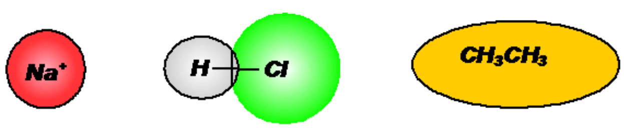
Forces

- Intra-molecular forces: Forces within the molecule. Generally stronger. Examples: Ionic, polar covalent and non-polar covalent bonds.
- Inter-molecular forces: Generally weaker. Examples: dipole-dipole and hydrogen bonding.

Where do we see the effects of intermolecular forces?

- They are the attraction that holds water into its liquid and solid shape.
- They are the forces that give water it's surface tension.
- They are the forces that break when going from a solid to a liquid or a liquid to a gas.

The type of molecule/ion determines the type of intermolecular force.

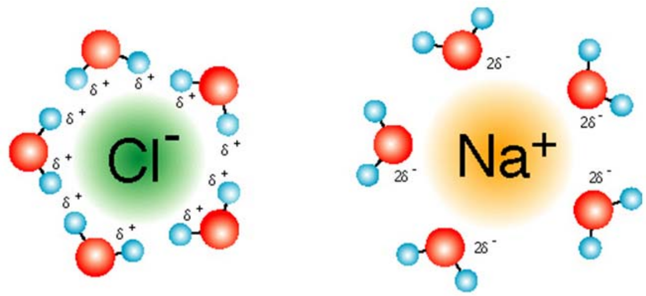


Ion | Molecules with **dipoles** | Non-polar molecules where the dipole has to be **induced**.

Types of intermolecular forces

- Ion-dipole (between ions and polar molecules)
- dipole-dipole (between 2 polar molecules)
- dipole-induced dipole (between a polar molecule and a non-polar molecules)
- induced dipole - induced dipole (between 2 non-polar molecules)

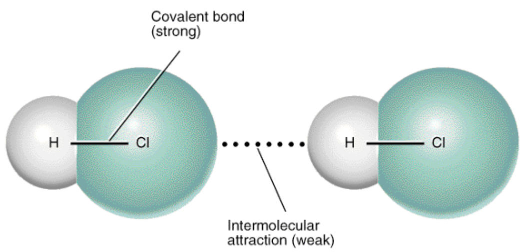
Water uses ion-dipole forces to dissolve salts



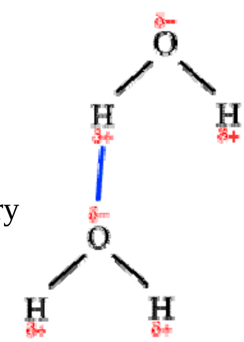
Slightly positive hydrogen are attracted to chlorine anions

Slightly negative oxygen are attracted to sodium cations

HCl uses Dipole-Dipole Forces



Hydrogen Bonding

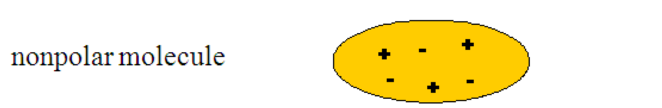


- A particularly strong example of dipole-dipole.
- Occurs when there is an OH, NH or FH bond. H is very small and O N and F are very electronegative.

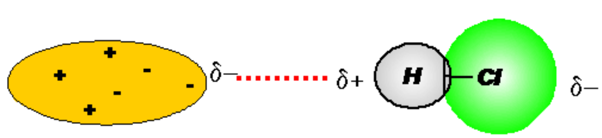
Conditions for Hydrogen Bonding

- Two important conditions must be met for hydrogen bonding to occur:
  - One molecule has a hydrogen atom attached by a covalent bond to an atom of N, O, or F.
  - The other molecule has an N, O, or F atom.

Dipole-induced dipole

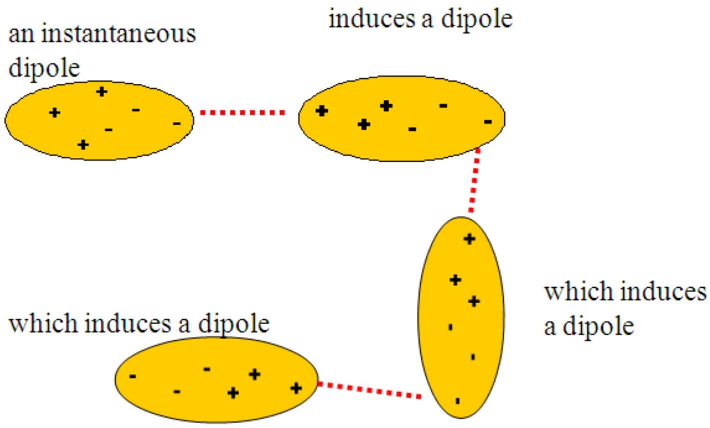


The dipole of the HCl attracts the electrons which induces a dipole in the non-polar CH_3 CH_3



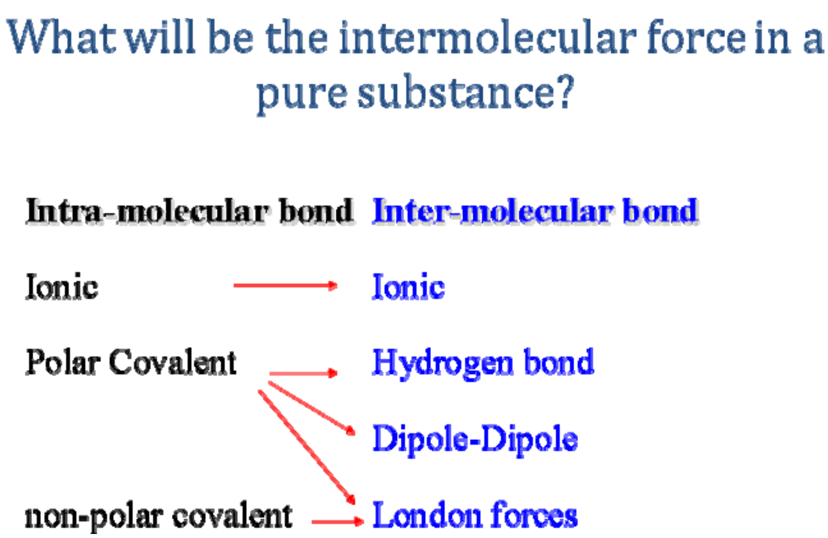
induced dipole - induced dipole

(also called London Forces)



Larger molecules have more attractions

What will be the intermolecular force in a pure substance?

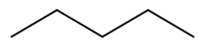


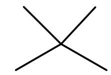
Factors that change boiling points

- Type of Intermolecular Forces
  - The stronger the force, the higher the b.p.
- Size
  - The larger the molecule, the more surface area, the higher the b.p.
- Shape
  - Branching decreases surface area which decreases b.p. (For organic compounds)

Boiling points

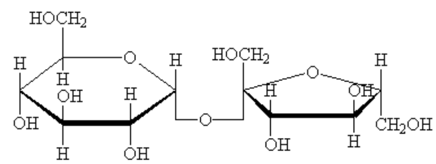
- Water, H_2O. 100 ^oC
- Methane, CH_4, -164 ^oC
- Pentane, C_5H_{12}, 36.2 ^oC
- 2,2-Dimethyl propane, 9.5 ^oC



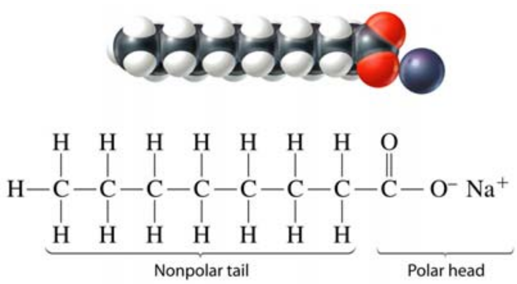


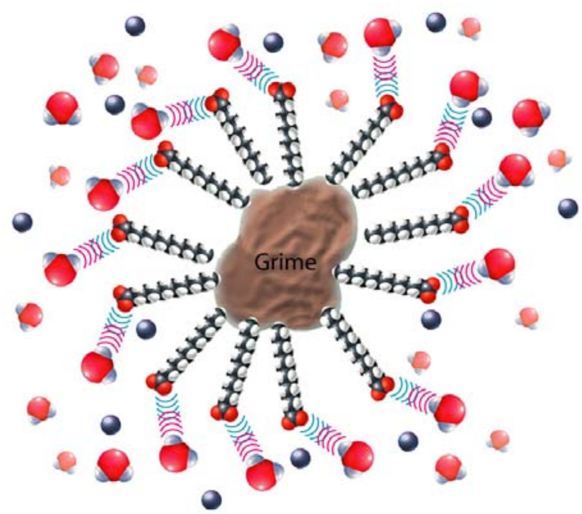
“Like Dissolves Like”

- Sugar (polar) dissolves in water (polar) but not in hexane (non-polar).
- Table salt, NaCl (very polar) dissolves in water (polar) but not in hexane (non-polar).
- Nanthalene, $C_{10}H_8$ (non-polar) dissolves in hexane (non-polar) but not in water (polar).



SOAP





What is happening to water during a phase change?

- When there are temperature changes, the molecules move faster.
- During melting some of H-bonds break.
- During vaporization, all the rest of the H-bonds break.
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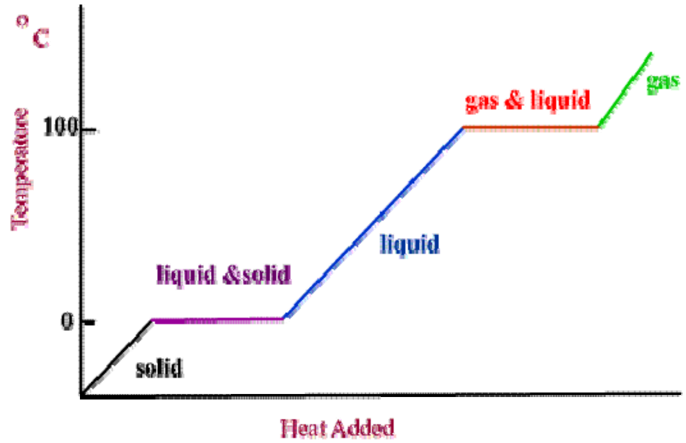
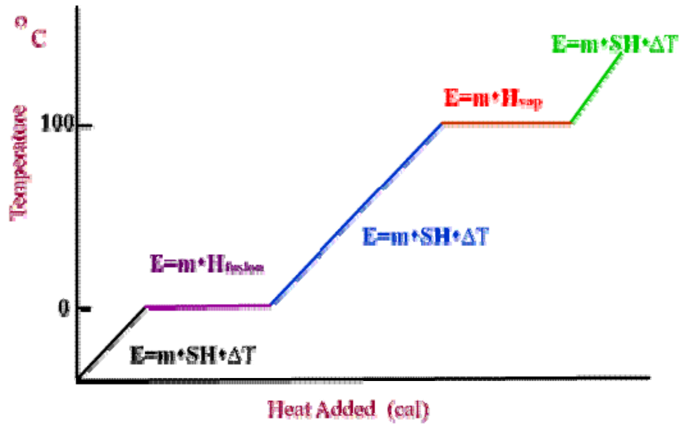
Names for Phase Changes

- Solid to Liquid: Melting
- Liquid to Gas: Boiling
- Solid to Gas: Sublimation
- Gas to Solid: Deposition
- Gas to Liquid: Condensation
- Liquid to Solid: Freezing

Heating curve



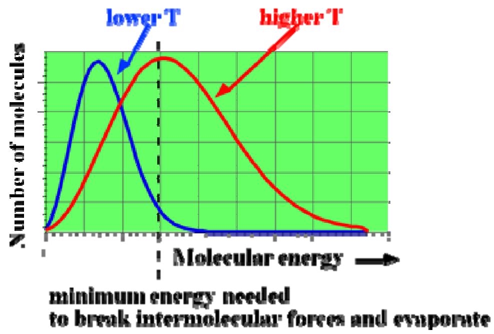


How much heat is required to vaporize 100. g of water from 37°C?

Why do Liquids Evaporate?

- There is a distribution of velocities among molecules, some move faster (higher energy), others slower (lower energy).
- To evaporate, molecules must have sufficient energy to break the intermolecular forces holding them.
- When the molecules break the intermolecular forces holding them, and move from the liquid phase into the gaseous phase, this is called EVAPORATION!

EVAPORATION



Equilibrium

- If the container is left open, all the molecules will eventually break free of the liquid and escape.
- If the container is closed, the molecules cannot diffuse away.
- Eventually some of the gaseous molecules collide and are forced back down into the liquid.
- *EQUILIBRIUM* is when the rate of molecules leaving the liquid phase is equal to the rate of molecules entering the liquid Equilibrium phase.

VAPOR PRESSURE

- When molecules escape from a liquid it is called evaporation. When they return to the liquid it is called liquefaction.
- At equilibrium, the partial pressure of vapor above the surface of the liquid can be measured. This is called the *VAPOR PRESSURE* of the liquid.
- A liquid has the same vapor pressure regardless of the surrounding gas or the pressure above it. The vapor pressure for that liquid is the same even if there is a vacuum above it.

More on Vapor Pressure

- As the temperature increases, the vapor pressure increases.
- Eventually the liquid boils when the vapor pressure is the same as the atmospheric pressure.
- The boiling point at 1 atm is called the normal boiling point.
- 100 ^{o}C is the normal boiling point for water because that's when it boils at 1atm.

Vapor Pressure

Liquid boils when its vapor pressure equals atmospheric pressure.

Temperature (°C)	Vapor pressure (mm Hg)	Temperature (°C)	Vapor pressure (mm Hg)
0	4.6	50	92.5
10	9.2	60	149.4
20	17.5	70	233.7
25[a]	23.8	80	355.1
30	31.8	90	525.8
37[b]	37.1	100	760.0
40	55.3		

[a]Room temperature
[b]Body temperature

Boiling point of water at various locations that differ in elevation

Location	Feet above sea level	P_{atm} (mm Hg)	Boiling point (°C)
top of Mt. Everest, Tibet	29,028	240	70
top of Mt. McKinley, Alaska	20,320	340	79
Leadville, Colorado	10,150	430	89
Salt Lake City, Utah	4,390	650	96
Madison, Wisconsin	900	730	99
New York, New York	10	760	100
Death Valley, California	−282	770	100.4

Surface Tension

- **Molecules on the surface are more attracted to the liquid molecules below them than to the gas molecules above.**
- **As a result, there is a pull towards the center of the liquid crowding the molecules at the surface.**
- **This pull is called *__SURFACE__* *__TENSION__* and establishes a layer that is hard to penetrate.**