Ch11

Condensed Matter

What makes solids and liquids stay together

… and keeps everything from being a gas.

version 1.6

FINAL EXAM

- Your Final Exam is Mon, May 22nd: **2:10-4:40 pm**
  you must take the final exam to pass the course.

- The exam will be in our Lab Room (Room 305 of this building)

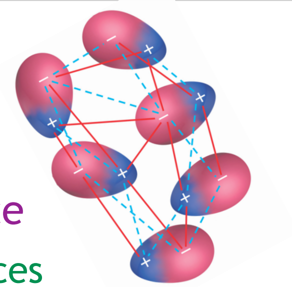
- If you want to get your final exam back,
  turn in a self addressed stamped (3 stamps)
  envelope with your final exam.

- Final Grades will be posted in WebSmart (not WebAccess!) after 2 weeks.
  - They are not available before then.

Ch11

Condensed Matter

inTRA — Within a molecule
inTER — Between molecules



- Solids & Liquids
  - Sticky Molecules
  - Intermolecular Force
    - Dipole Dipole Forces
      - Scales with polarity
    - London Forces
      - Scales with molecular size & shape
    - Hydrogen Bonding (O, N, F)
- Properties of Condensed Matter
  - IMF Controls Many Properties
  - When IMF & Energy Balance
    - Fluidity, Viscosity
  - Cohesion & Adhesion
    - Surface Tension & Droplet Shape
    - Capillary Action
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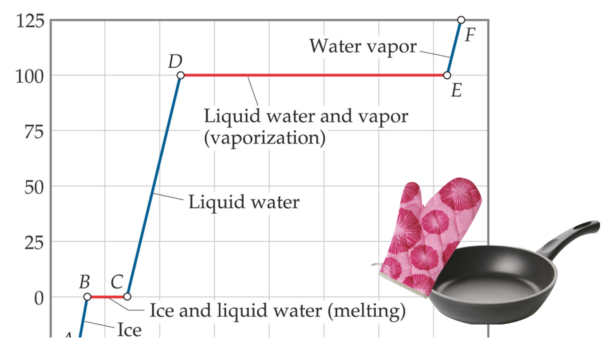
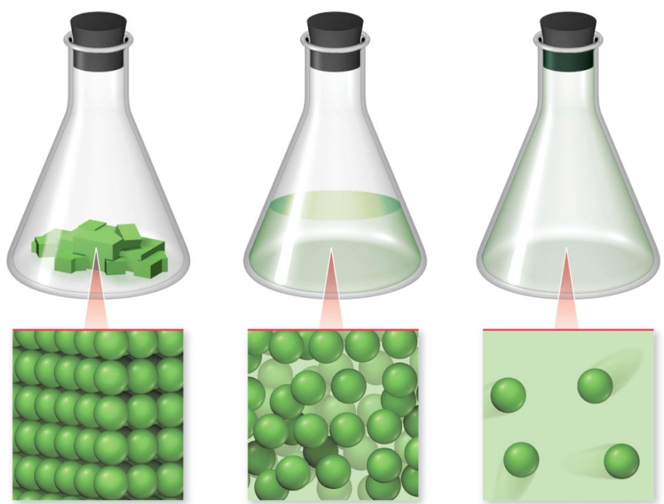
- Characteristic Temperatures
  - Vaporization/Condensation
    - at the Boiling Point [ΔH_{vap}]
  - Critical Temperature
  - Sublimation/Fusion
    - at the Melting Point [ΔH_{fus}]
- Heat & Matter
  - Connecting the pieces…
    - T < mp (heating solids)
    - T = mp (melting solids)
    - bp > T > mp (heating liquids)
    - T = bp (boiling liquids)
    - T > bp (heating gases)



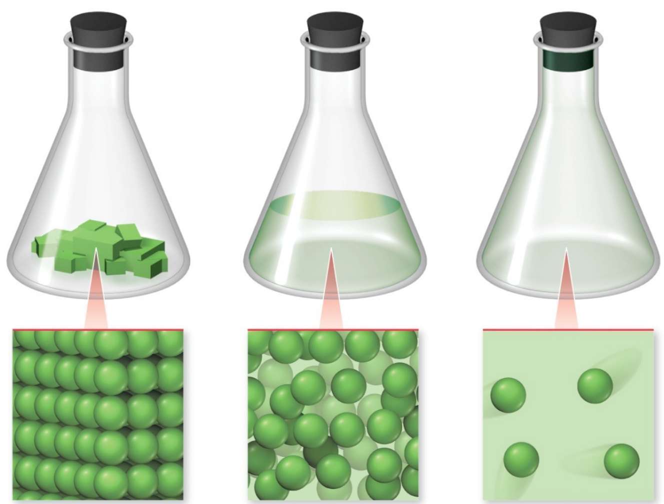
Three States of Matter

- Matter can exist three different ways.
  - It can tightly packed (solid)
  - It can be wadded up randomly (liquid)
  - It can be stretched thin (gas)
- It's still the same matter, just in a different state.
- Like your favorite shirt can be neatly folded, wadded up, or stretched over a coat hanger.
- It still the same shirt, but being in a different state means it may have some different properties.



Three States of Matter

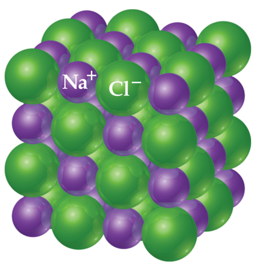
	properties	Gas	Liquid	Solid
Bulk	**Shape**	Variable	Variable	Fixed
Bulk	**Volume**	Variable	Fixed	Fixed
	Compressible	Extremely	Slight	None
Particles	**Structure**	Flexible	Flexible	Fixed
Particles	**Density**	Least	Compact	Most
Particles	**Cohesion**	Least	Some	Most
	Energy	Most	Some	Least



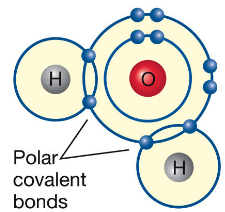
Atomic Forces

- Atomic forces are what keeps all matter from falling apart into a gas of atoms.
- We've been looking at the forces <u>within</u> molecules.
  - **Intra**molecular forces.
  - These forces are strong

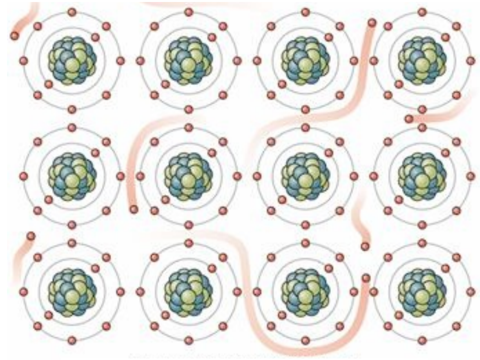
inTRA — Within a molecule
inTER — Between molecules



400-4000 kJ/mol



150-1100 kJ/mol



75-1000 kJ/mol

- Now let's look at the forces <u>between</u> particles.
  - **Inter**particular and **Inter**molecular forces
  - These are medium and weak forces.

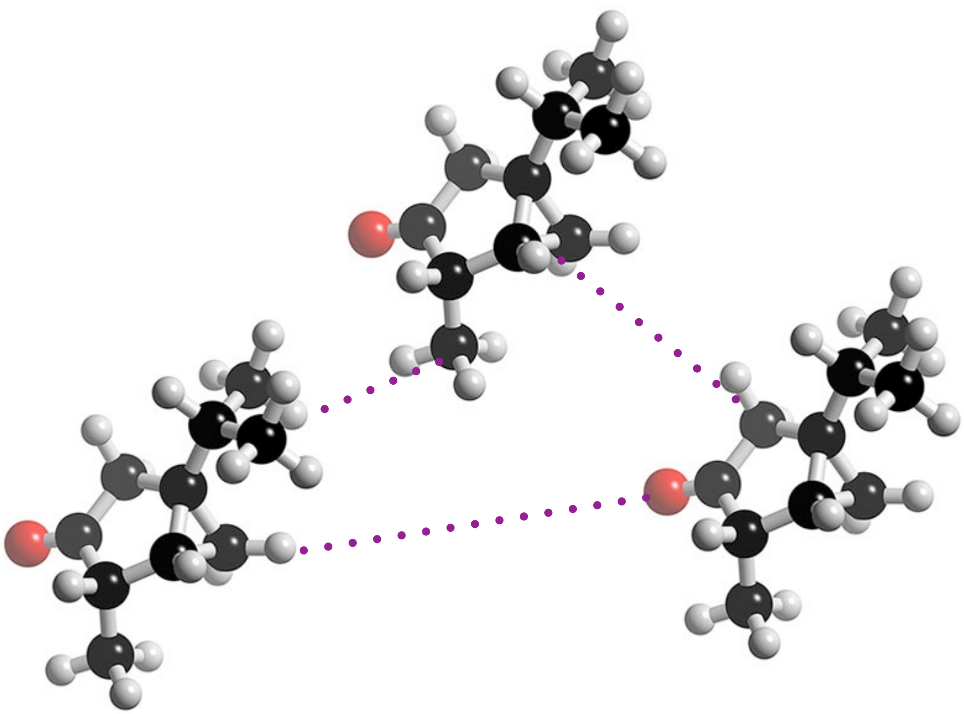
Intermolecular Forces

inTRA — Within a molecule
inTER — Between molecules

Intramolecular Forces
Interparticular Forces
Intermolecular Forces

We have already talked about these. (Intramolecular Forces, Interparticular Forces)

This chapter we will talk about these. (Intermolecular Forces)

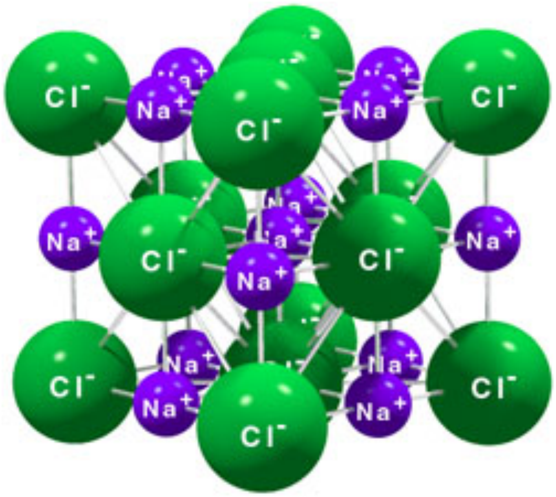


They are the forces
between molecules.

Interparticular Forces

- Interparticular Forces act between particles
  - Not necessarily molecules
  - The most significant inter particular force are ion-dipole forces.
  - We discussed ion-dipole forces in chapter 4.
  - Ion-dipole forces are medium strength.

40-600 kJ/mol



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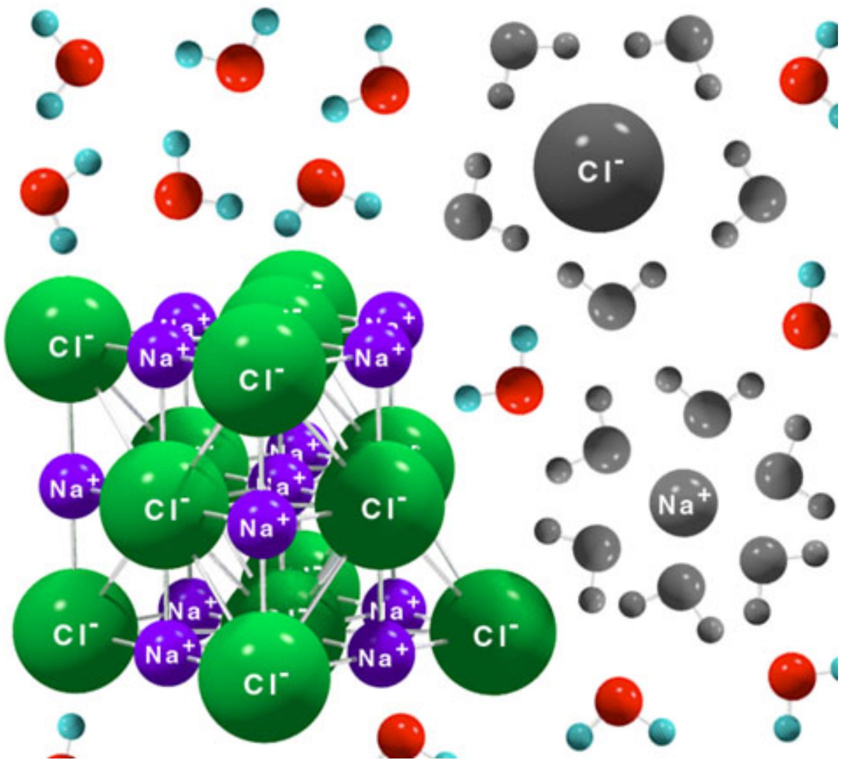
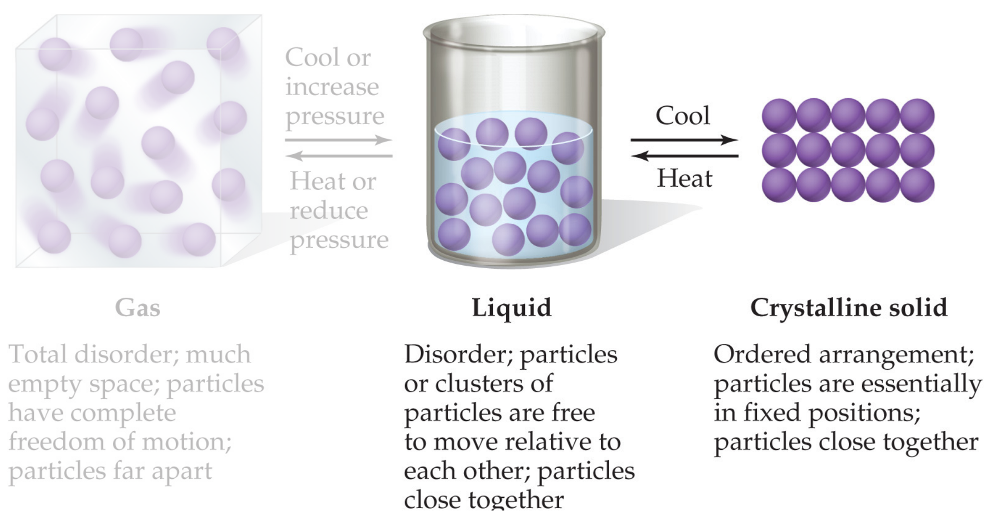


Intermolecular Forces Control States

- The fundamental difference between states of matter is the distance between particles.
- Because in the solid and liquid states particles are closer together, we refer to them as condensed phases.
- The state a substance is in at a particular temperature and pressure depends on two antagonistic entities:
  - the kinetic energy of the particles;
  - the strength of the attractions between the particles.

TABLE 11.1 ■ Some Characteristic Properties of the States of Matter

Gas	Assumes both the volume and shape of its container Is compressible Flows readily Diffusion within a gas occurs rapidly
Liquid	Assumes the shape of the portion of the container it occupies Does not expand to fill container Is virtually incompressible Flows readily Diffusion within a liquid occurs slowly
Solid	Retains its own shape and volume Is virtually incompressible Does not flow Diffusion within a solid occurs extremely slowly

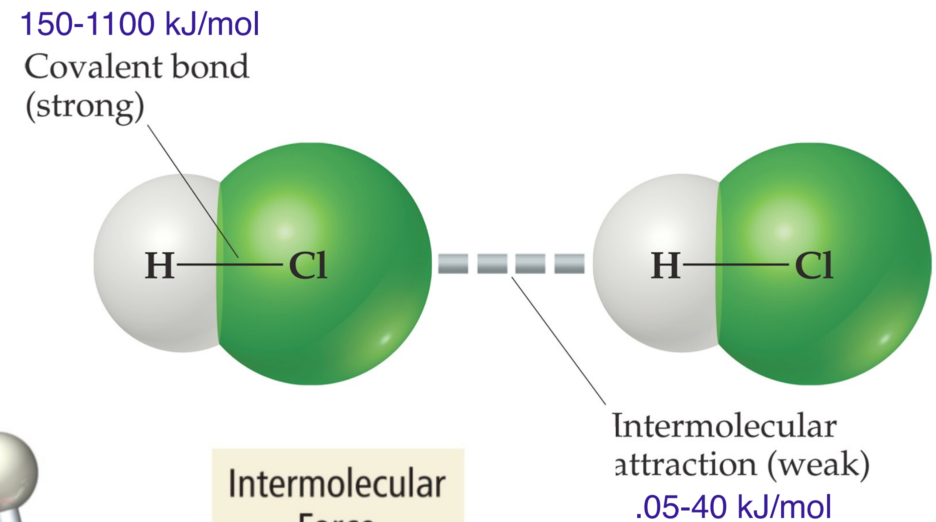


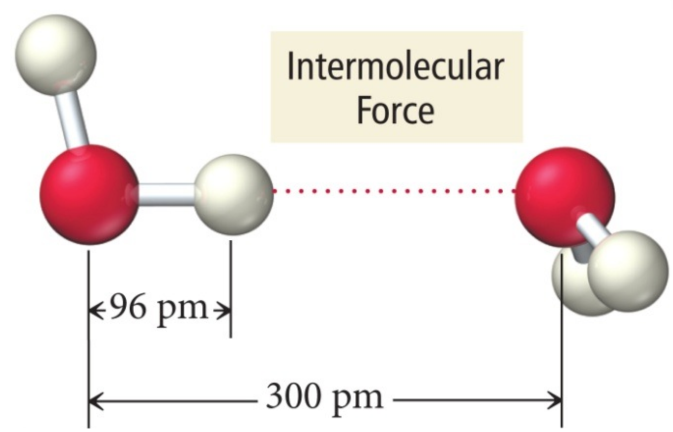
Intermolecular Forces

inTRA — Within a molecule
inTER — Between molecules

- Intermolecular forces are called Van der Walls forces.
- The attractions between molecules are not nearly as strong as the other forces.
- The forces are about 30x less than intermolecular forces.
  - These forces are small, but don't underestimate it's importance.
- Intermolecular forces control many important physical properties:
  - boiling point
  - melting point
  - vapor pressure
  - viscosity
  - hardness
  - surface tension …

greater IF means higher boiling pint, more viscous liquid, harder solid…





"Size matters not … the [intermolecular] force is in everything."

- Yoda, 1980

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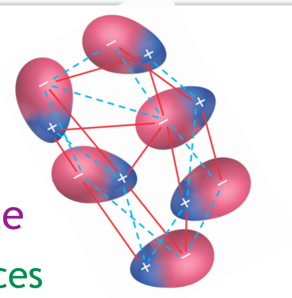



- Characteristic Temperatures
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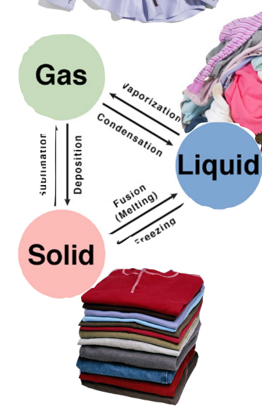
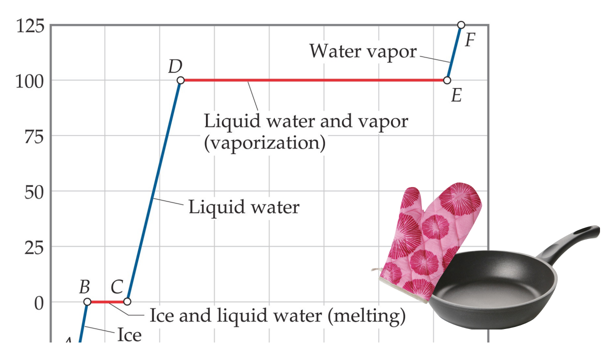




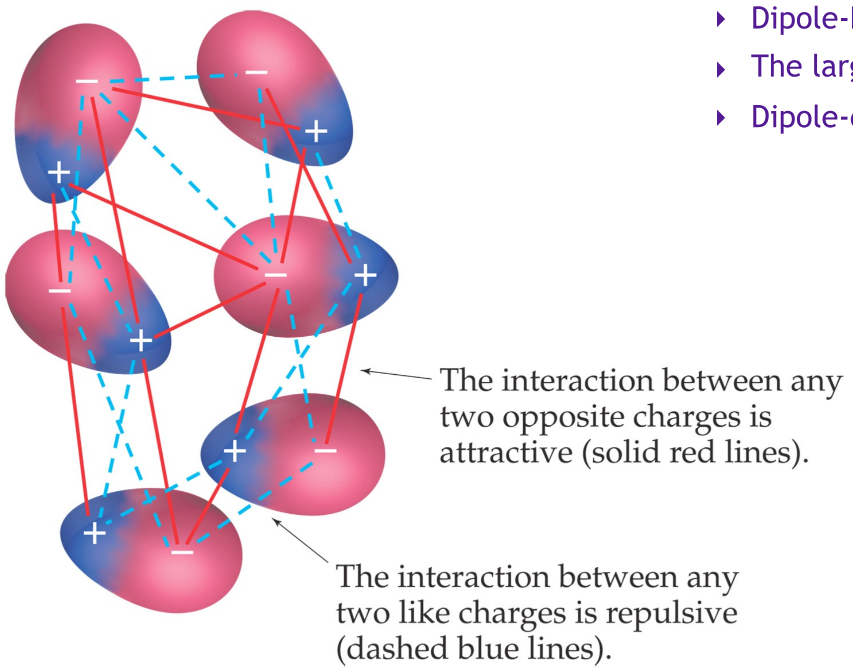
Dipole-Dipole Interactions

There are three intermolecular forces:

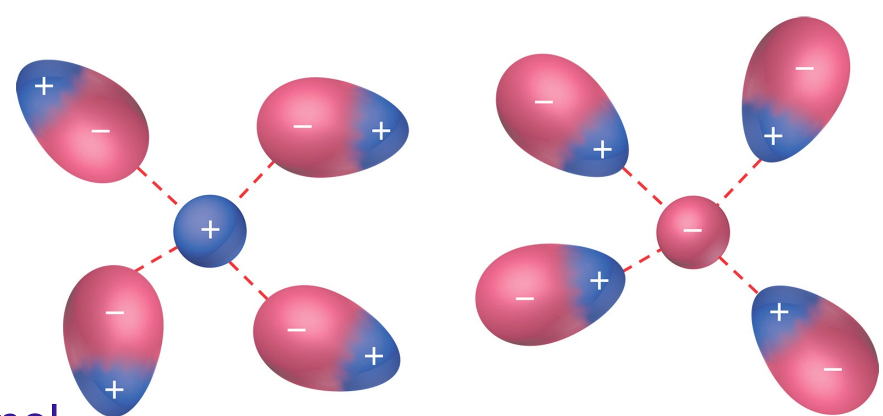
1. Dipole-dipole interactions
2. London-Dispersion forces
3. Hydrogen Bonding

- Ion-dipole interactions (an interparticular force), are important in solutions of ions.
- The strength of these forces are what make it possible for ionic substances to dissolve in polar solvents.
- Molecules that have permanent dipoles are attracted to each other in a similar way.
  - The positive end of one is attracted to the negative end of the other and vice-versa.
  - These forces are only important when the molecules are close to each other.
- Dipole-Dipole forces are weaker than ion-dipole forces.
- The force of attraction is about 5-25 kJ/mol.
- Dipole-Dipole forces increase with a molecules polarity.
- The larger the dipole moment, larger the dipole-dipole force.
- Dipole-dipole interactions exist *only in polar substances*.

5-25 kJ/mol



40-600 kJ/mol



Cation-dipole attractions

Anion-dipole attractions

Dipole-Dipole Interactions

There are three intermolecular forces:

1. Dipole-dipole interactions
2. London-Dispersion forces
3. Hydrogen Bonding

Substance	Molecular Weight (amu)	Dipole Moment μ (D)	Boiling Point (K)
Propane, $CH_3CH_2CH_3$	44	0.1	231
Dimethyl ether, CH_3OCH_3	46	1.3	248
Methyl chloride, CH_3Cl	50	1.9	249
Acetaldehyde, CH_3CHO	44	2.7	294
Acetonitrile, CH_3CN	41	3.9	355

Dipole-dipole interactions are apparent in trends like boiling point.

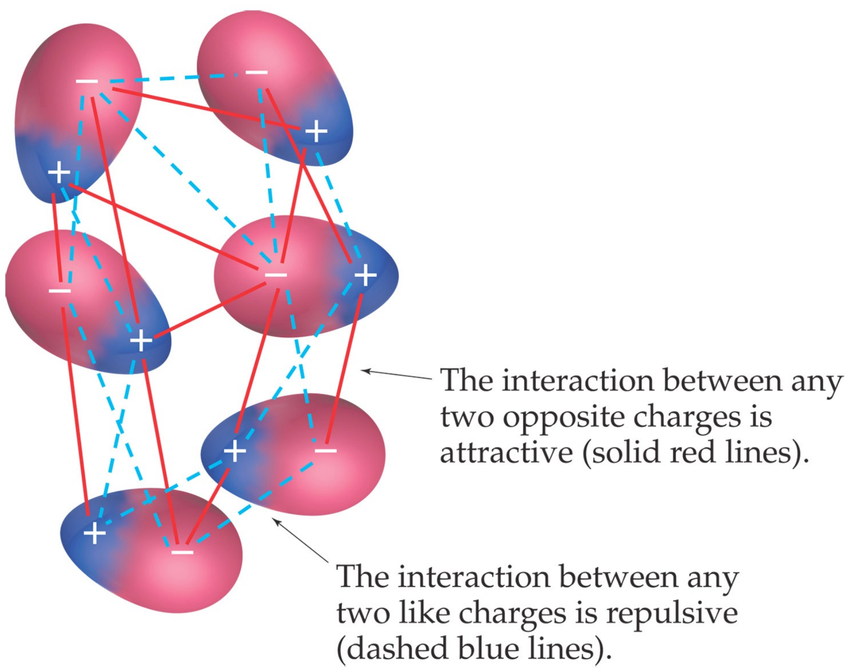
For similar molecules:

the more polar the molecule ⬆

higher intermolecular forces ⬆

the higher is its boiling point ⬆

5-25 kJ/mol



Dipole-dipole interactions exist only in polar substances!

London-Dispersion Forces

There are three intermolecular forces:

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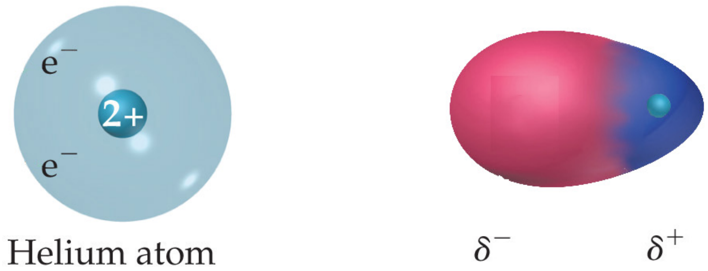
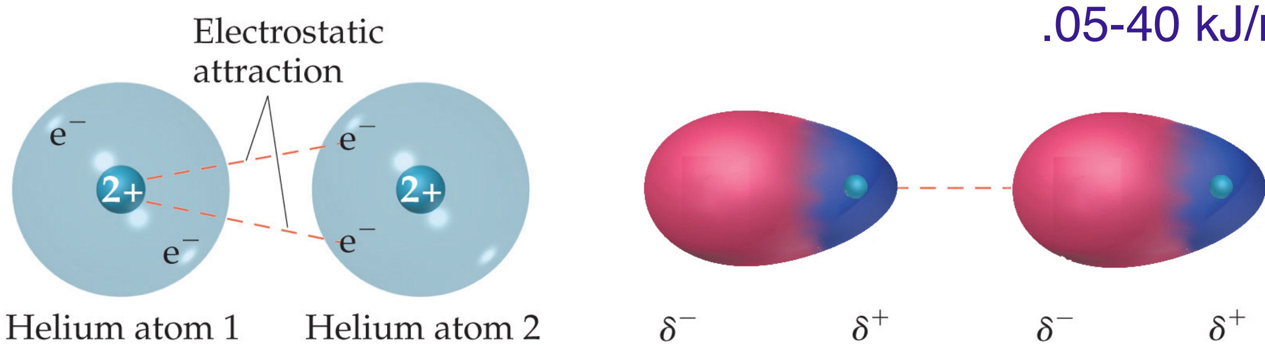


- While the electrons in the 1s orbital of helium would repel each other (and, therefore, tend to stay far away from each other), it does happen that they occasionally wind up on the same side of the atom.
- At that instant, then, the helium atom is polar, with an excess of electrons on the left side and a shortage on the right side.
- If it's close enough, adjacent helium atoms could be effected by that momentary dipole.
- The electrons on the left side of helium atom 2 repel the electrons in the cloud on helium atom 1.
- Even though the two atoms are not polar overall, the momentary shift in electrons causes them both become polar for a very short time — and that polar moment causes them to be attracted.
- London dispersion forces (or just dispersion forces), are attractions between an instantaneous dipole and an induced dipole.
- These forces are present in all molecules, whether they are polar or nonpolar.
- The tendency of an electron cloud to distort in this way is called polarizability.



.05-40 kJ/mol

London-Dispersion Forces

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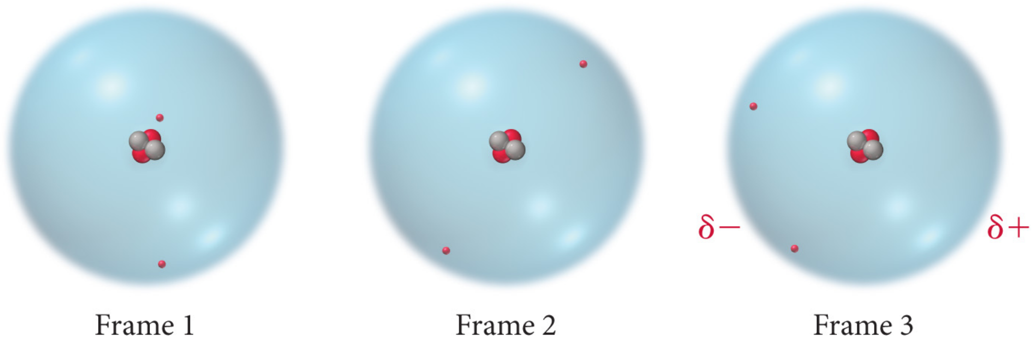
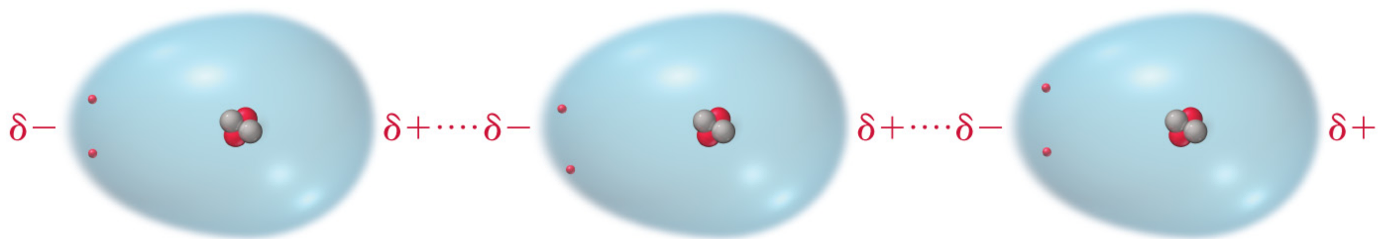
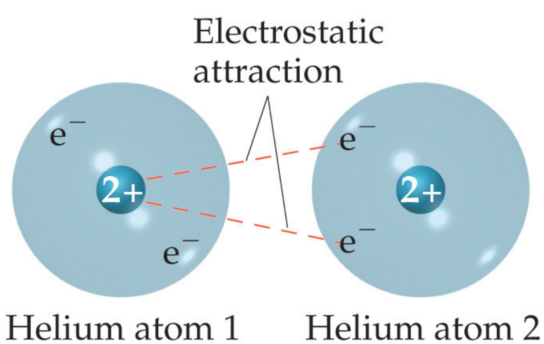
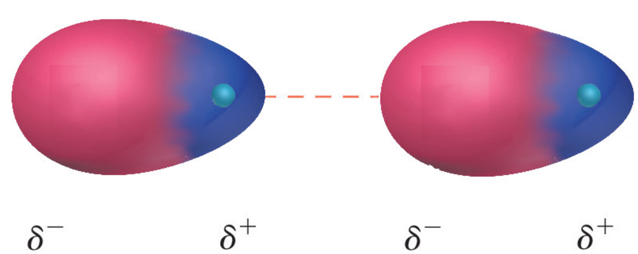






.05-40 kJ/mol



London-Dispersion Forces

There are three intermolecular forces:

1. Dipole-dipole interactions
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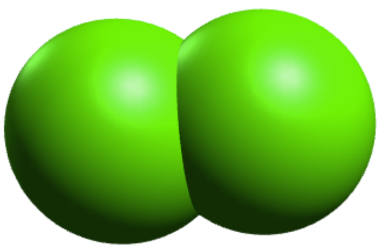
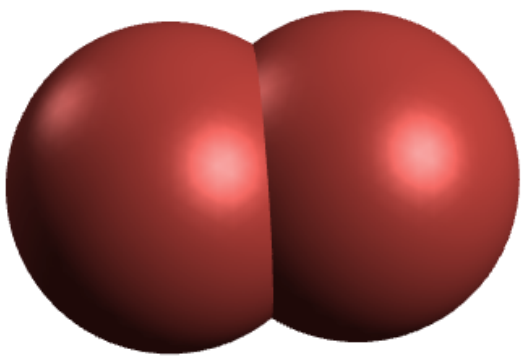
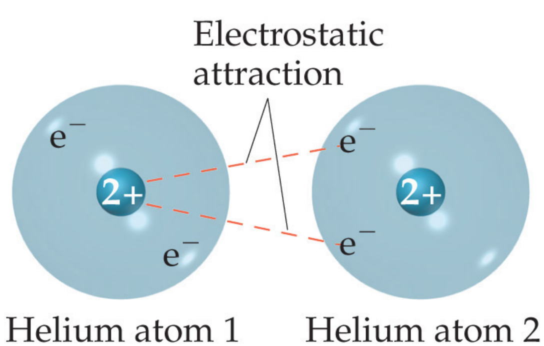
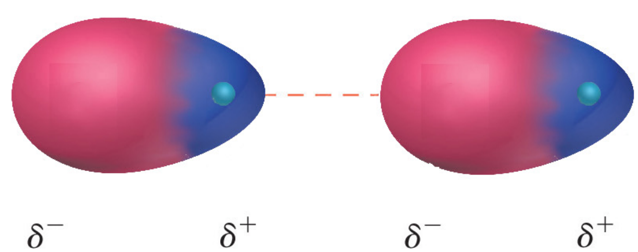
- Dipoles are relative to both the charge and the distance between them.
- Molecules with a large surface area can have greater distance between the momentary charges on their surface.
- So molecules with a greater surface area have greater dispersion forces.
- Br_2 has greater dispersion forces than Cl_2
- Br_2 therefore has a higher boiling point.
- Mass is usually relative to surface area, so you can usually say molecules with more mass will have higher dispersion forces.
- Therefore molecules with higher mass will be more viscous, higher boiling, etc.





Halogen	Molecular Weight (amu)	Boiling Point (K)	Noble Gas	Molecular Weight (amu)	Boiling Point (K)
F_2	38.0	85.1	He	4.0	4.6
Cl_2	71.0	238.6	Ne	20.2	27.3
Br_2	159.8	332.0	Ar	39.9	87.5
I_2	253.8	457.6	Kr	83.8	120.9
			Xe	131.3	166.1





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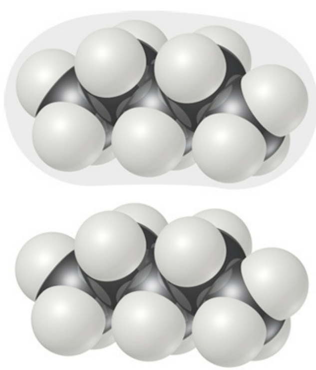
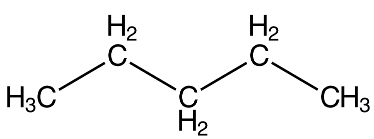
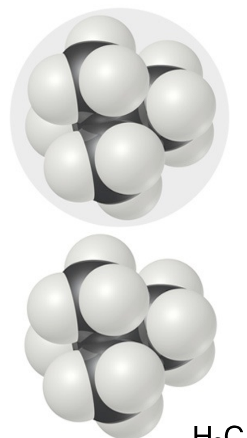
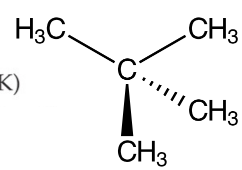


n-Pentane
(bp = 309.4 K)



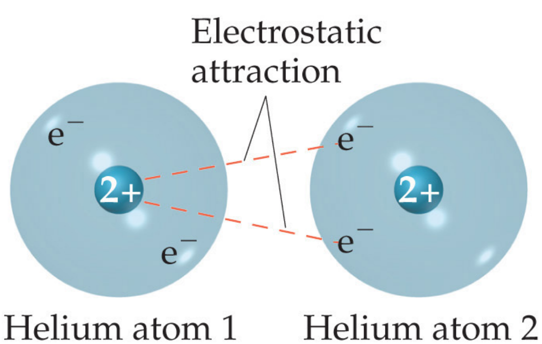


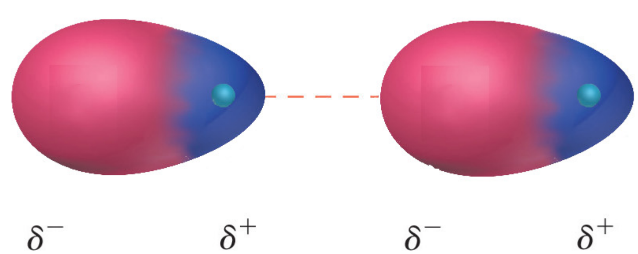
Neopentane
(bp = 282.7 K)



- Shape is also important.
- Two molecules with the same mass may have different surface areas depending on their shape.
- For example molecules with more branches produce more compact shapes.
- They have less surface area, therefore less dispersion forces.
- And lower boiling points.

.05-40 kJ/mol





Hydrogen Bonding

There are three intermolecular forces:

1. Dipole-dipole interactions
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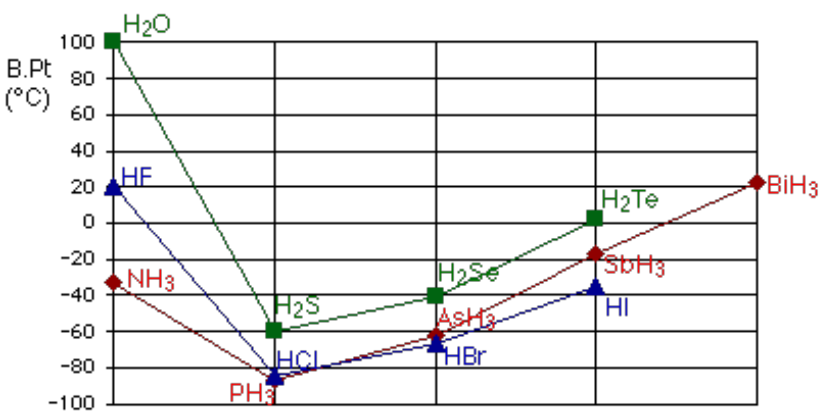
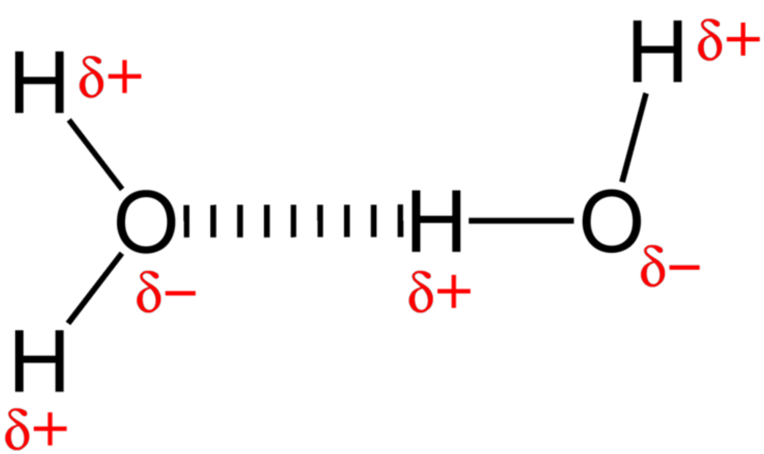
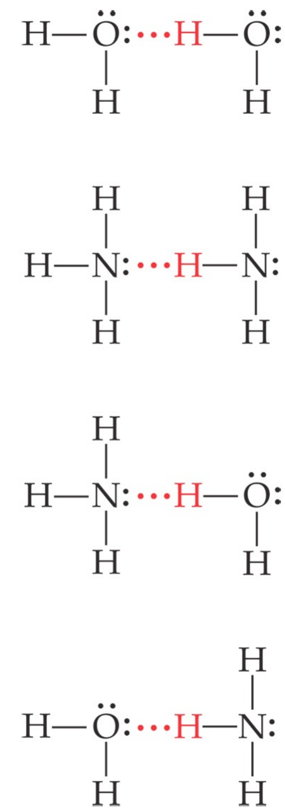
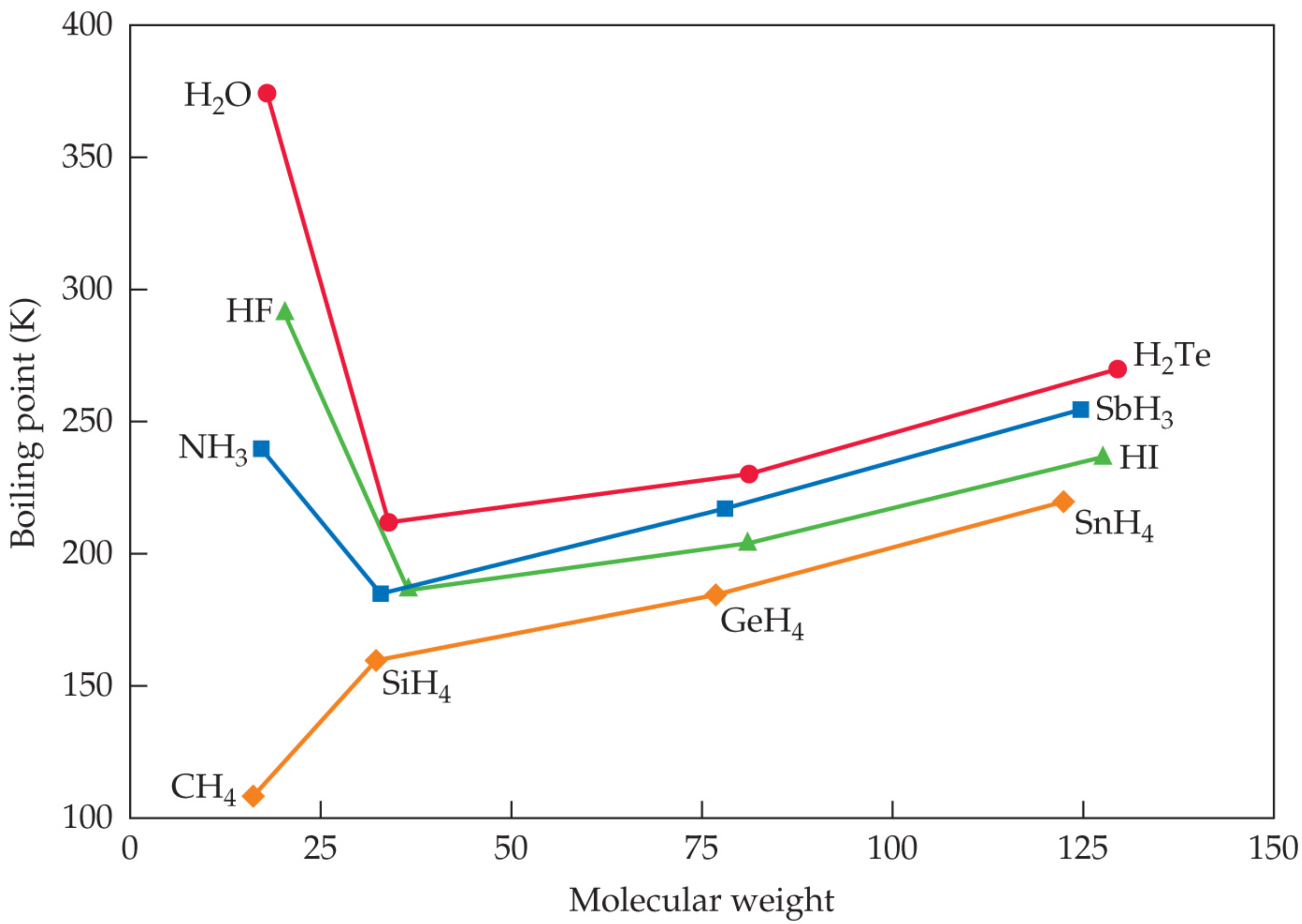




- Whenever a molecule has a bond between H and F, O, or N we see very strong IF forces.
- F, O, and N are the most electronegative elements.
- They pull so much electron density away from the hydrogen that it leaves it almost a bare proton.
- Then we get electrostatic at action between the exposed hydrogen and F, O, and N in adjacent molecules.



Hydrogen Bonding



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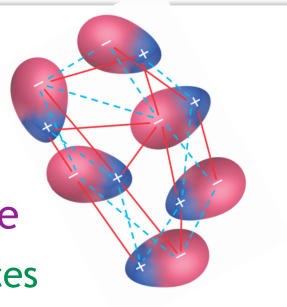



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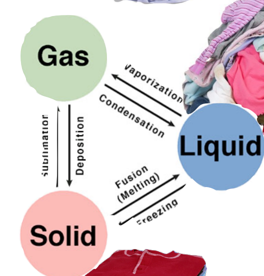


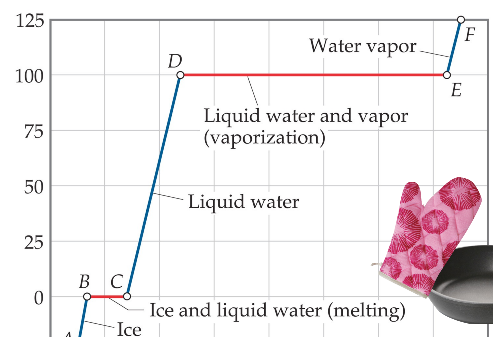

Hydrogen Bonding

There are three intermolecular forces:

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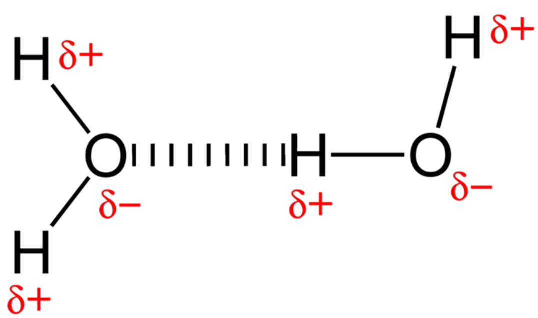
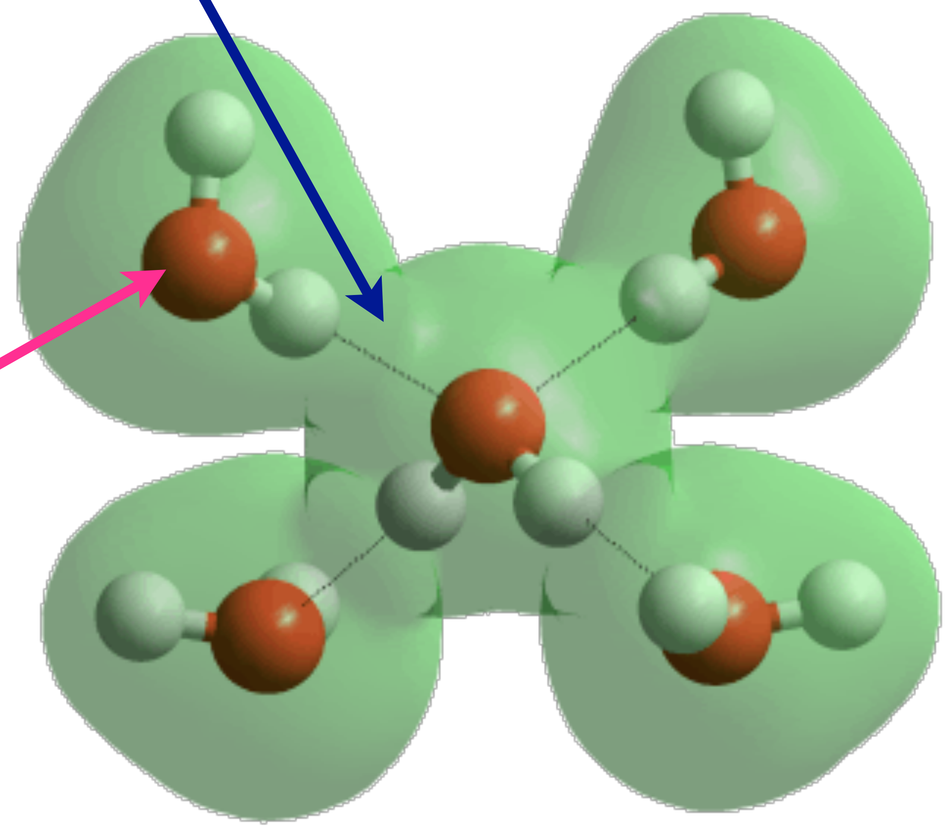
O,N & F are very electronegative
so they have strong
negative charge for the other molecules
hydrogen to be attracted to.



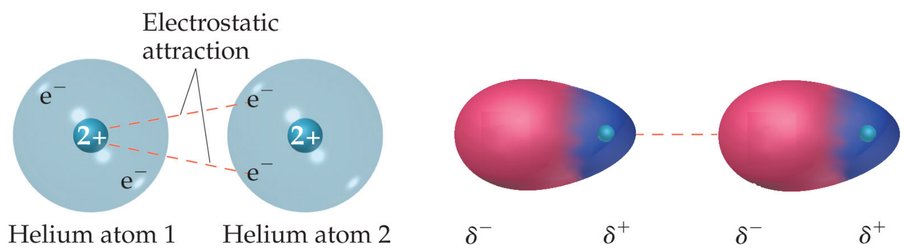
10-40 kJ/mol

That same high electronegatively causes
them to strip their own hydrogens of
electrons, exposing the positively charged
nucleous. Making their hydrogens more
hungry for negative charge.

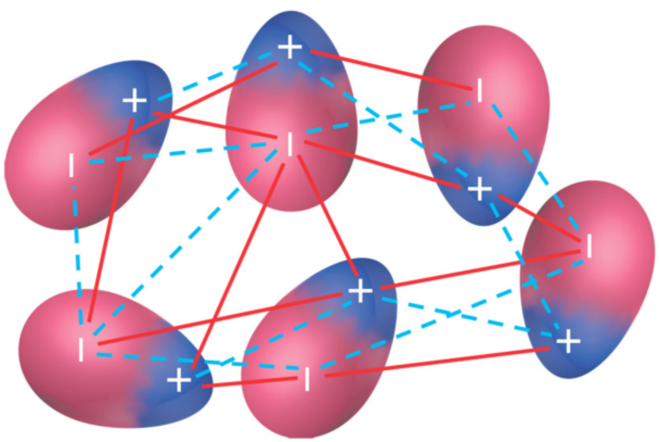


Hydrogen Bonding

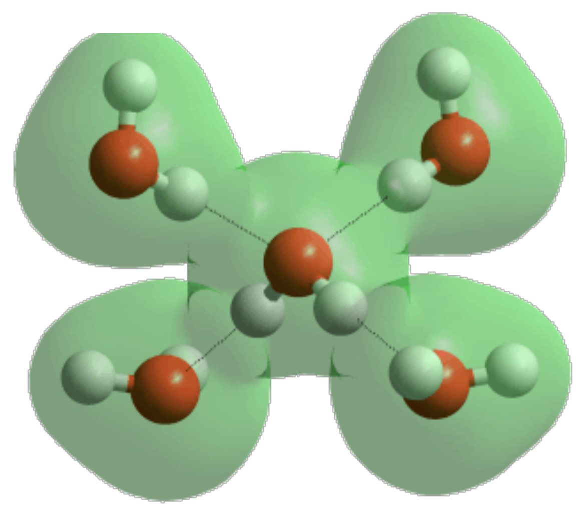
- If hydrogen bonding exists in a molecule, it will be the most important factor in considering IF.



Dispersion
in all molecules
05-40 kJ/mol



Dipole-Dipole
only in Polar Molecules
5-25 kJ/mol

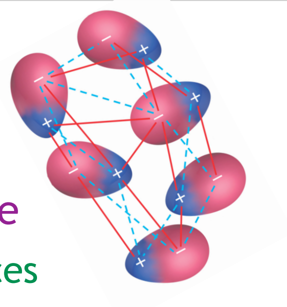


Hydrogen Bonding
only in molecules with OH, NH, or FH
10-40 kJ/mol

Ch11

Condensed Matter

inTRA — Within a molecule
inTER — Between molecules



- Solids & Liquids
  - Sticky Molecules
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    - Dipole Dipole Forces
      - Scales with polarity
    - London Forces
      - Scales with molecular size & shape
    - Hydrogen Bonding (O, N, F)
- Properties of Condensed Matter
  - IMF Controls Many Properties
  - When IMF & Energy Balance
    - Fluidity, Viscosity
  - Cohesion & Adhesion
    - Surface Tension & Droplet Shape
    - Capillary Action
    - Meniscus Shape

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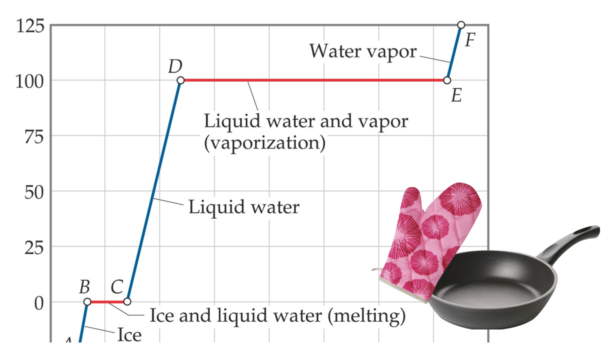
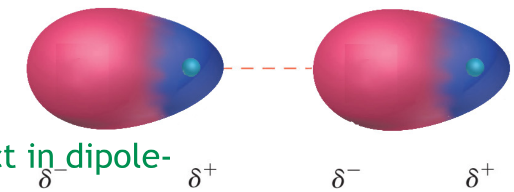
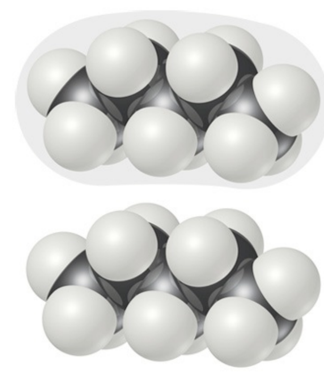


Finding which molecule has greater IF

- All molecules have dispersion forces, but only some have dipole-dipole forces. A polar molecule will usually have greater IF than a non polar one.
- If both molecules are polar:
  - Differences in EN effect IF.
    - Polarity can be influenced by differences in electronegativity.
    - Small changes in ΔEN have big impact on dipoles, therefore big impact in dipole-dipole forces.
  - Symmetry can effect IF.
    - Polarity can be influenced by shape.
    - Symmetric molecules with have less polarity than asymmetric ones.

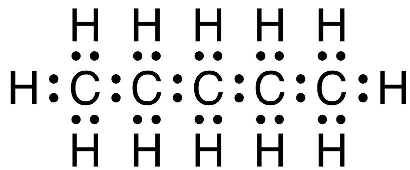
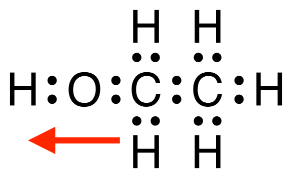
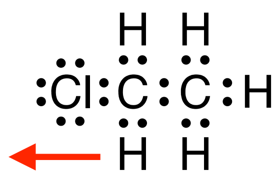
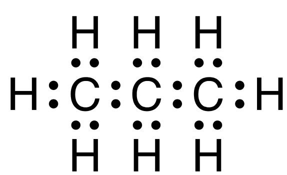


- If both molecules are non-polar:
  - Mass can be an indicator of IF.
    - Greater mass generally corresponds to greater surface area, therefore greater dispersion forces.
  - Branched molecules have lower IF.
    - More branched molecules tend to be more compact, therefore lesser surface area and lesser dispersion forces.



- Differences in IF can indicated differences in boiling point, viscosity, hardness, and other properties of matter.

Order by decreasing boiling point…

			H-Bond	Dipole	London
3	C_5H_{12}		✗	✗	✓ Bigger
1	C_2H_5OH		✓	✓	✓
2	C_2H_5Cl		✗	✓	✓
4	C_3H_8		✗	✗	✓ Smaller

Ch11

Condensed Matter

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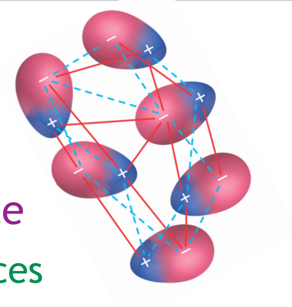

- Properties of Condensed Matter
  - IMF Controls Many Properties
  - When IMF & Energy Balance
    - Fluidity, Viscosity
  - Cohesion & Adhesion
    - Surface Tension & Droplet Shape
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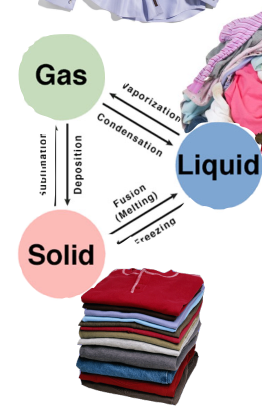
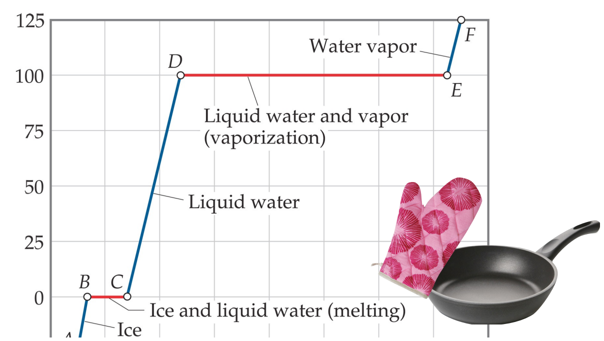
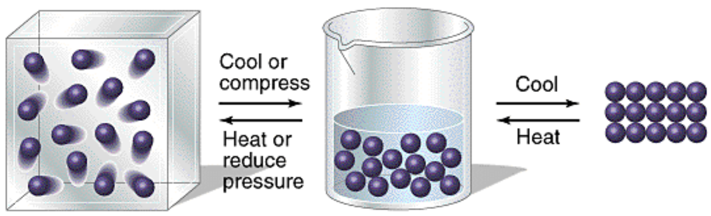




Liquid is the Between State

- In gases kinetic energy is much greater than IF, so IF forces have little effect.
- In solids, IF forces dominate kinetic energy locking atoms into rigid structures.
- In liquids, we're at a place between those two extremes.
  - In liquids there is an equilibrium between these two competing forces tha produces some unique physical properties.

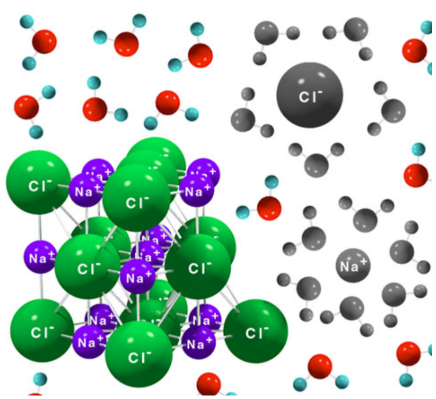


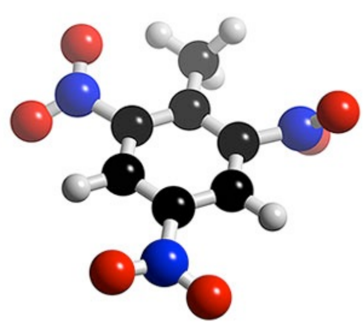
	Gas	Liquid	Solid
Definite Volume	No	Yes	Yes
Rigid Shape	No	No	Yes
Compressibility	High	Small	None
Particle Separation	Large	Between	Small
Particle Energy/Speed	High	Between	Low
Particle Attraction	Weak	Between	Strong

Intermolecular Forces

inTRA — Within a molecule
inTER — Between molecules

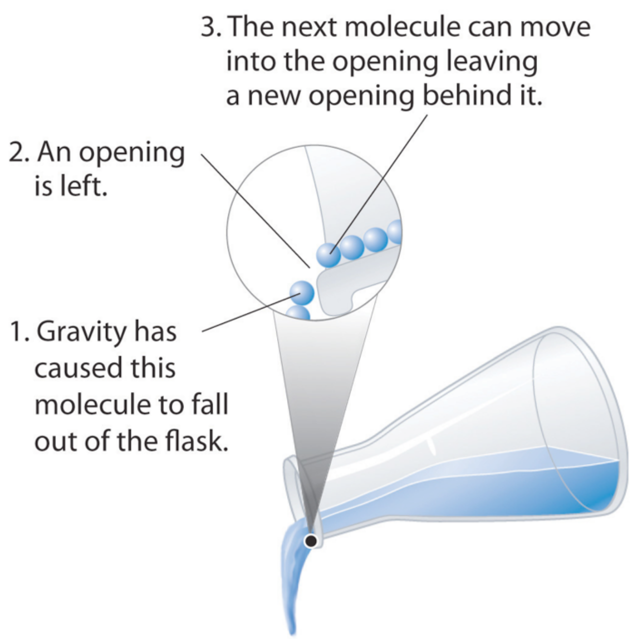
- Intramolecular forces verses Intermolecular forces
  - Strengths (intramolecular forces are stronger)
  - Differences (intermolecular forces are everywhere)
- The three states of matter
  - The gas phase has little or no intermolecular forces
  - The condensed states of matter have significant intermolecular forces
    - Their properties are defined by intermolecular forces
- Intermolecular Forces
  - Dipole-dipole forces
    - Only in polar molecules
    - Magnitude relative to Polarity
  - Dispersion forces
    - In all molecules
    - Magnitude relative to Polarizability (not the same thing as polarity!)
      - Look at molecular size (mass can indicate) and shape
  - Hydrogen Bonding
    - Only in OH, NH, and FH molecules
    - Magnitude relative to number of OH, NH, and HF groups
- Physical Properties of Different States
  - What makes liquids unique
- Unique Properties of Liquids
  - Cohesion & Adhesion
    - Fluidity
    - Viscosity
    - Surface Tension & Droplet Shape
    - Capillary Action
    - Meniscus Shape





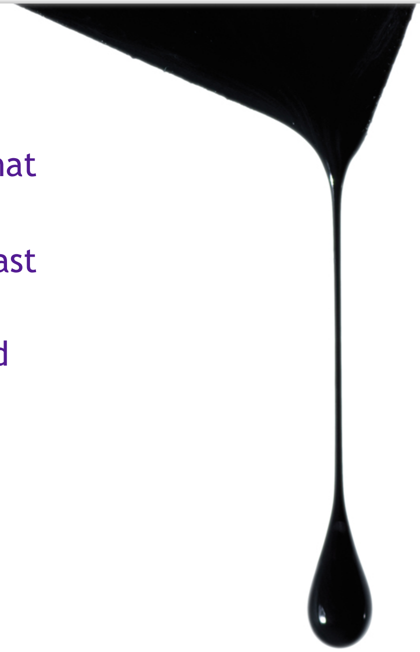
Fluidity

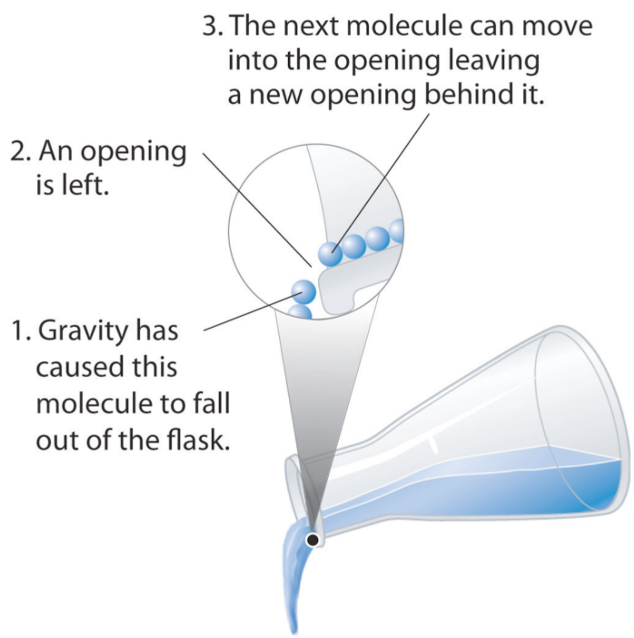
- Molecules in a liquid are in constant motion.
- When the flask is tilted, molecules move to the left and down due to the force of gravity.
- The result is liquid flowing out of the container.
- Fluidity is the capacity of a substance to flow.
- Liquids are characterized by their tendency to flow.
- Resistance of a liquid to flow is viscosity.
- Viscosity is the attraction between molecules that holds the molecules in position. Resisting flow.



Viscosity

- Resistance of a liquid to flow is viscosity.
- Viscosity results from the attraction between molecules that holds the molecules in position. Resisting flow.
- It is related to the ease with which molecules can move past each other.
- Viscosity increases with stronger intermolecular forces and decreases with higher temperature.
- Viscosity is not density, it's about the force not the mass.
- The length of a oil drop measures viscosity by competing intermolecular forces with the force of gravity.





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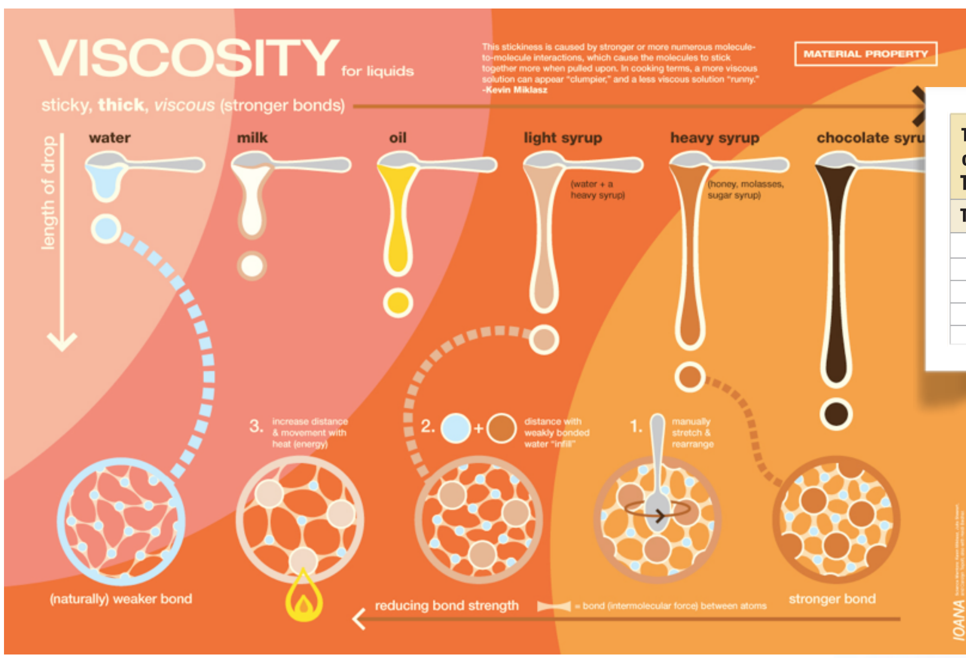

TABLE 11.6 Viscosity of Liquid Water at Several Temperatures

Temperature (°C)	Viscosity (cP)
20	1.002
40	0.653
60	0.467
80	0.355
100	0.282

Ch11

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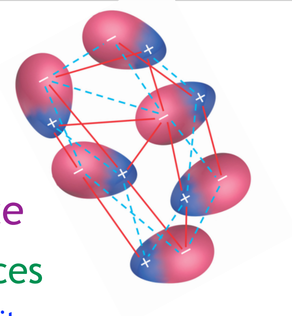

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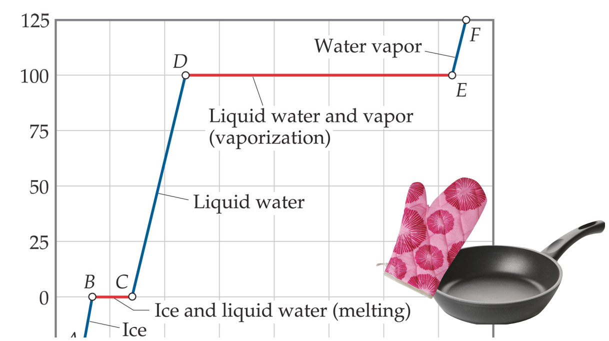


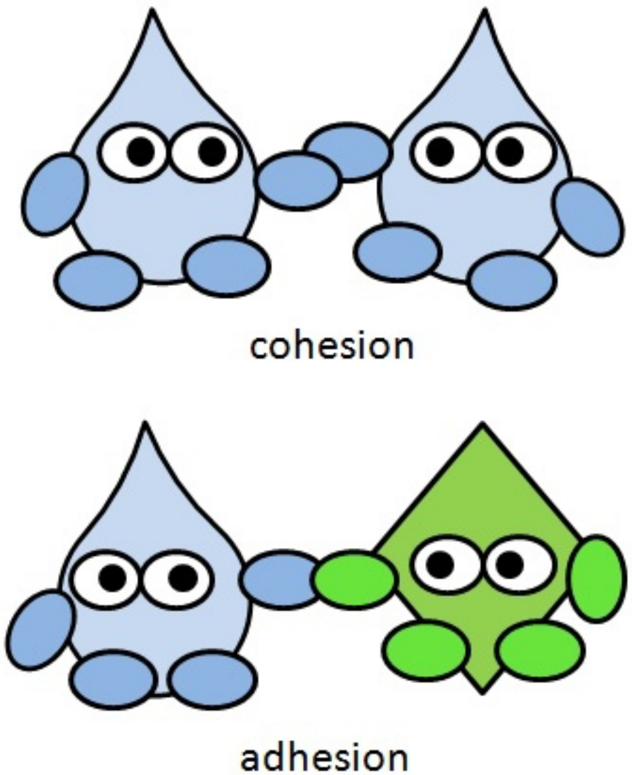
Water Dripping from a Leaf

- Intermolecular forces hold the droplets together.
- They also make the water stick to the leaf.

Water Dripping from a Leaf

- ✓ Cohesion is the attraction between molecules within the substance.
- ✓ Adhesion is the attraction between the molecules of one substance and a second substance.

Water Dripping from a Leaf



- ✓Cohesion is the attraction between molecules within the substance.
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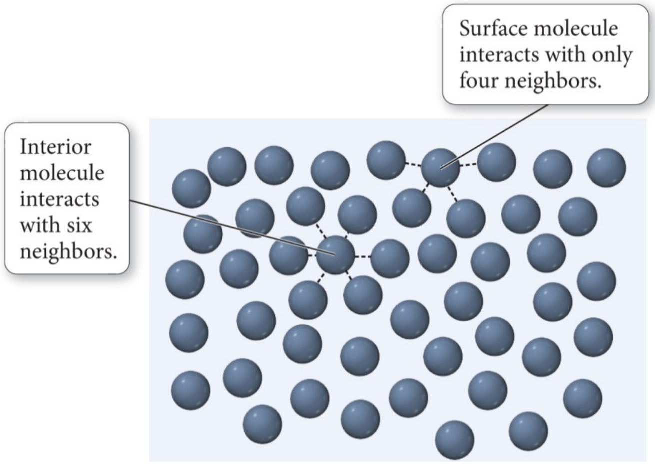
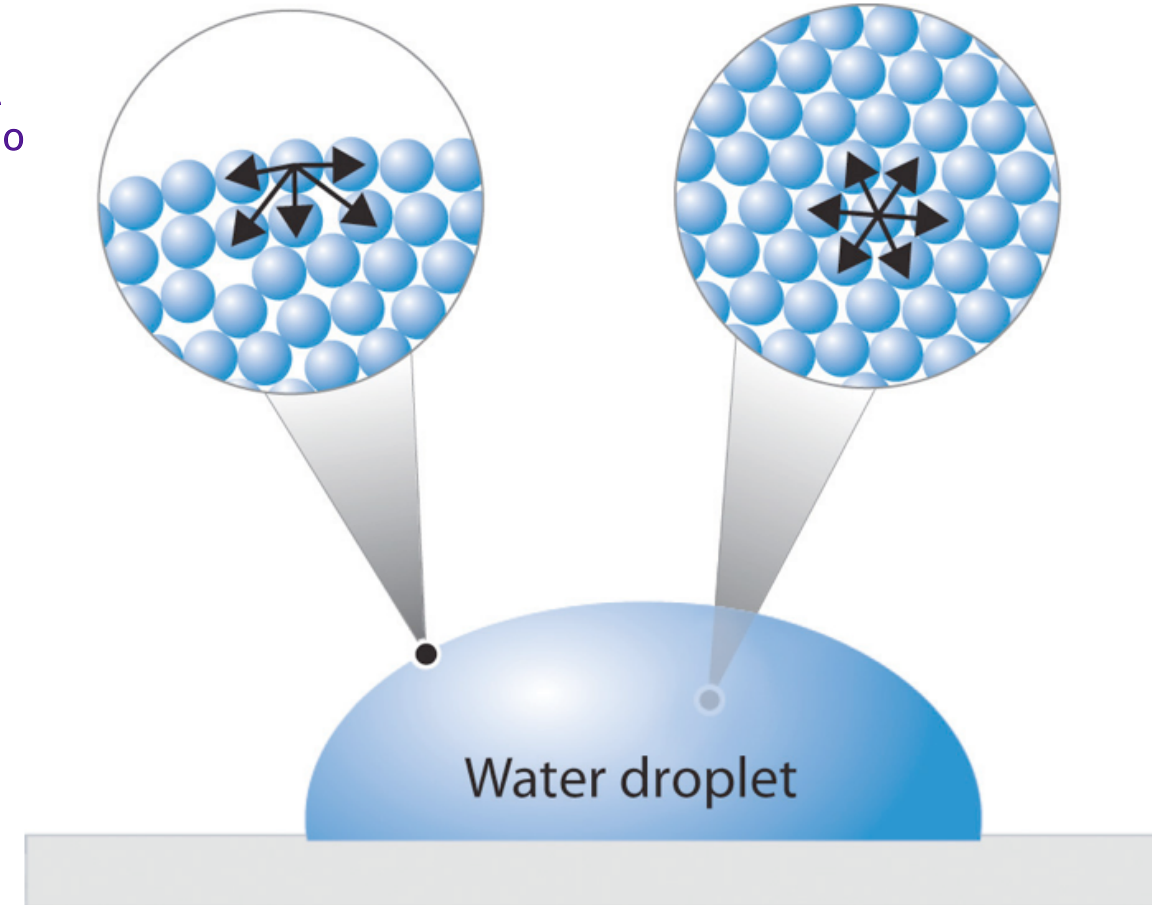
Cohesive Forces

- Cohesive forces pull the molecules on the surface of a liquid inwards.
- The surface molecules also pull each other towards themselves.
  - Like links in a net.
- The result is liquids minimize their surface area.
- It becomes harder to separate molecules at the surface of a liquid, than at the interior o the liquid.
- Surface tension is the property of liquids to minimize their surface area.





Cohesive Forces

- A sphere has the least surface area per volume of any geometric shape.
- So liquids prefer to form spheres.
- Spheres can become flattened when the force of gravity competes with it's cohesive forces.

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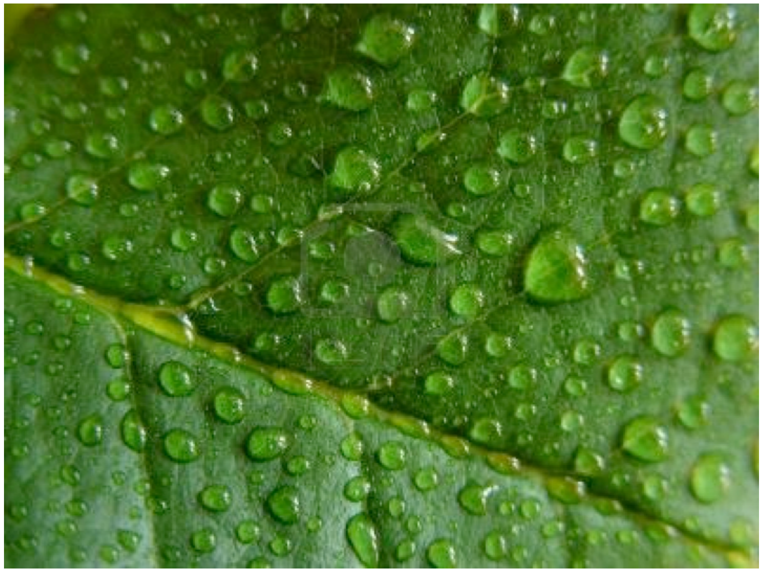

Cohesive Forces

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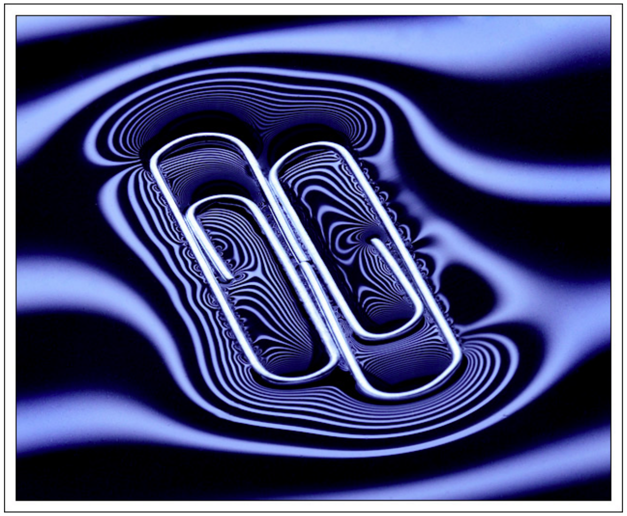
- So liquids prefer to form spheres.
- Spheres can become flattened when the force of gravity competes with it’s cohesive forces.
- If the cohesive forces are strong enough they can support considerable weight.



Ch11

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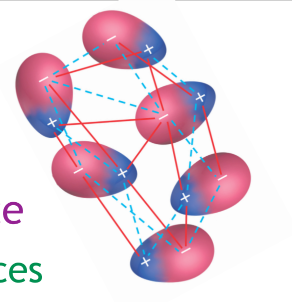

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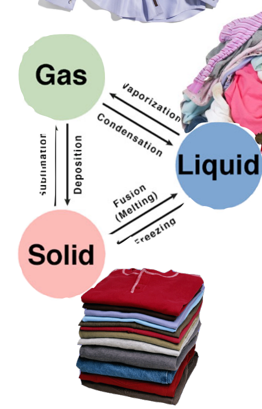
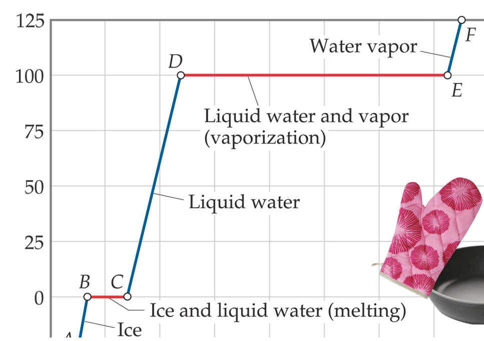
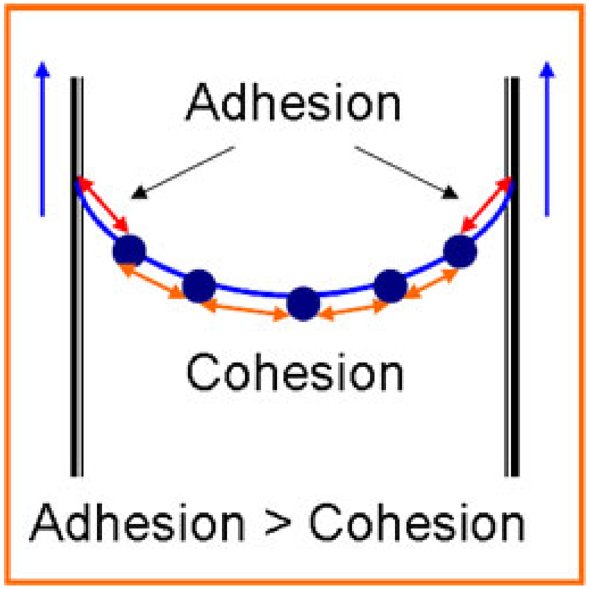




Capillary action is the spontaneous rising of a liquid in a narrow tube.



Capillary action results from the cohesive forces within the liquid and the adhesive forces between the liquid and the walls.

If the adhesive forces between the liquid and the walls of its container exceed the cohesive forces between the molecules of the liquid, the liquid will climb the walls of the container.

When a liquid is placed in a glass cylinder, the surface of the liquid shows a curve called the meniscus.

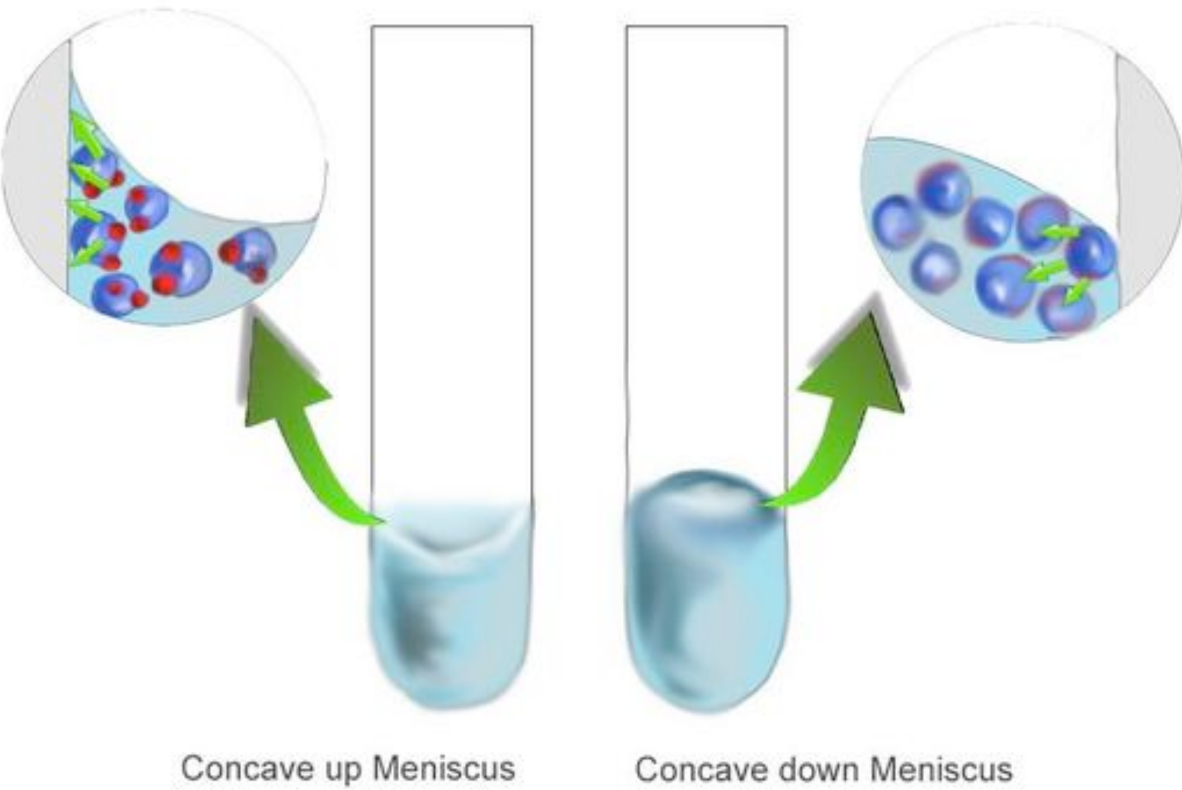
If adhesive forces are stronger than cohesive forces, the meniscus is concave

(example: water is more strongly attracted to glass than itself).

If adhesive forces are weaker than cohesive forces, the meniscus is convex

(example: mercury is more strongly attracted to itself than glass).

When a liquid is placed in a glass cylinder, the surface of the liquid shows a curve called the meniscus.



Ch11

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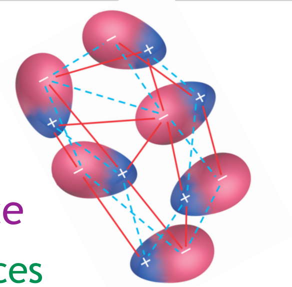

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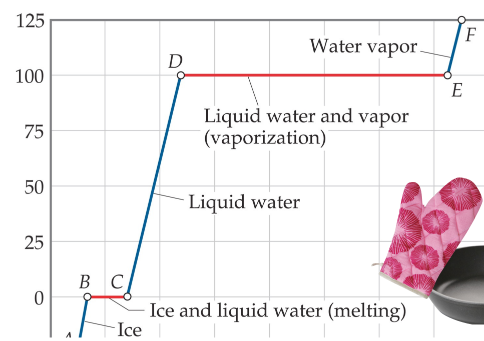
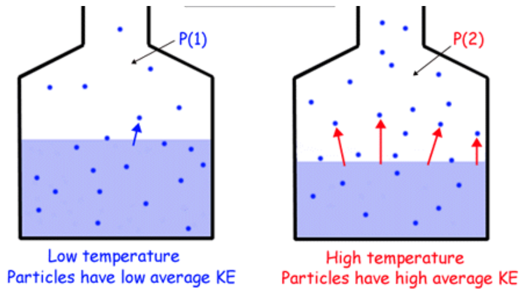
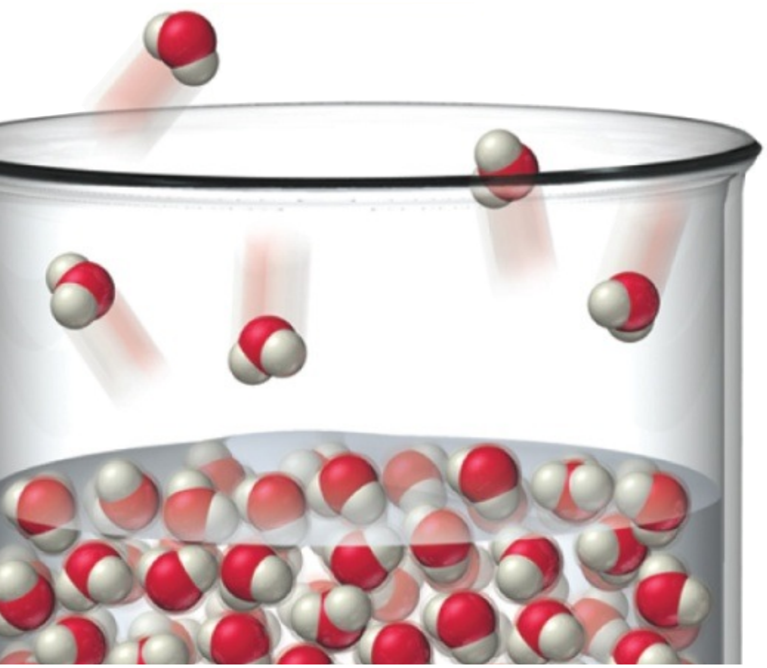
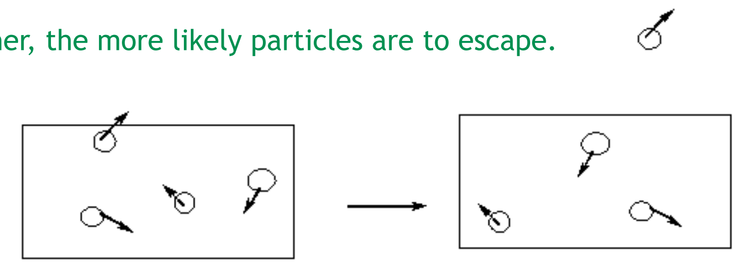


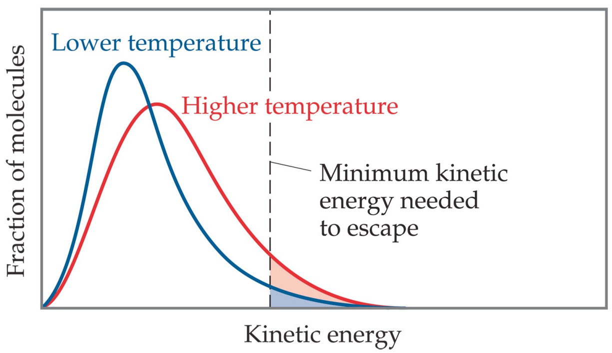
Vaporization

- Vaporization is the process where thermal energy (motion) overcomes intermolecular forces to change the state of matter from liquid to gas.
  - Vaporization is temperature dependent.
    - At any temperature some molecules in a liquid have enough energy to escape.
    - When those particles reach the surface, they overcome the cohesive forces of the liquid and fly free as gas molecules.
    - As the temperature rises, the fraction of molecules that have high enough energy to escape increases.
  - Vaporization is dependent on the surface area.
    - The more spread out the substance is, the easier it is for a fast moving particle to find the edge and break out.
  - Vaporization depends on the intermolecular forces.
    - The weaker the particles hold each other, the more likely particles are to escape.



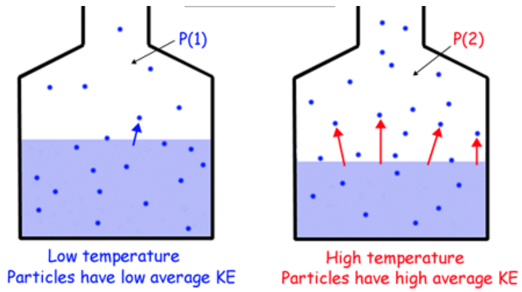


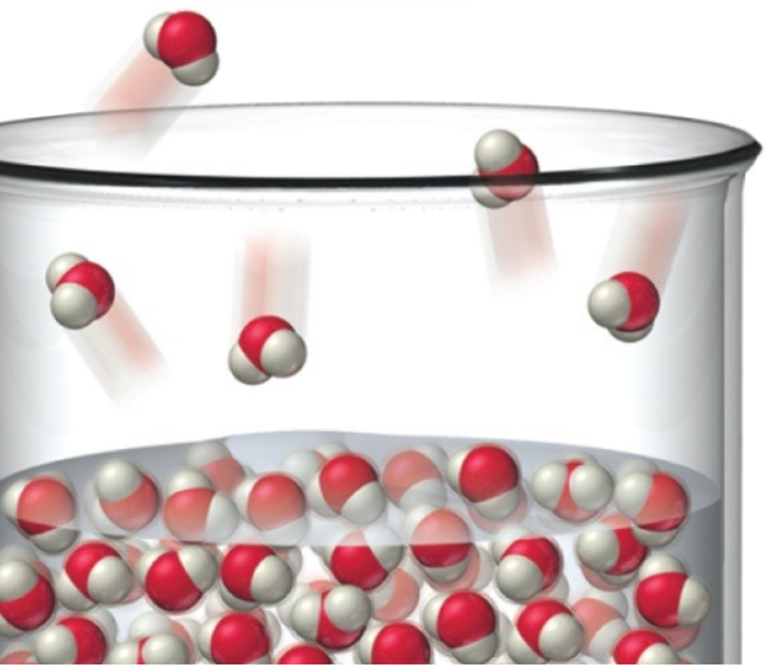


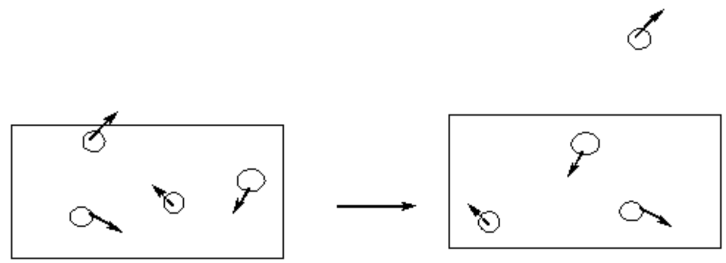


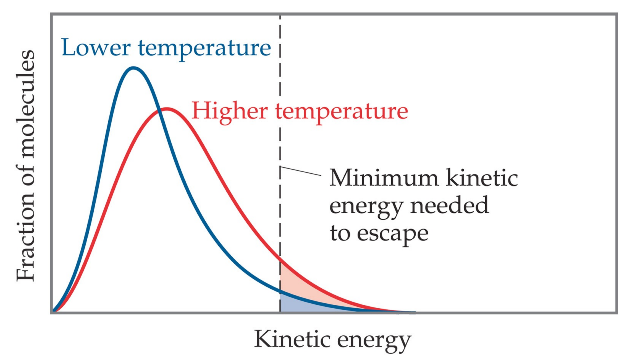
Vaporization

- Condensation is the process of a gas succumbing to intermolecular forces and becoming a liquid.
  - Condensation and Vaporization can both happen at the same temperature.
    - At any temperature some molecules in the gas have less thermal energy that the cohesive forces of the liquid.
    - When those particles strike the surface of the liquid, the cohesive forces cause the molecule to stick to the liquid.
    - As the temperature decreases, the fraction of molecule that have low enough energy to get trapped increases.



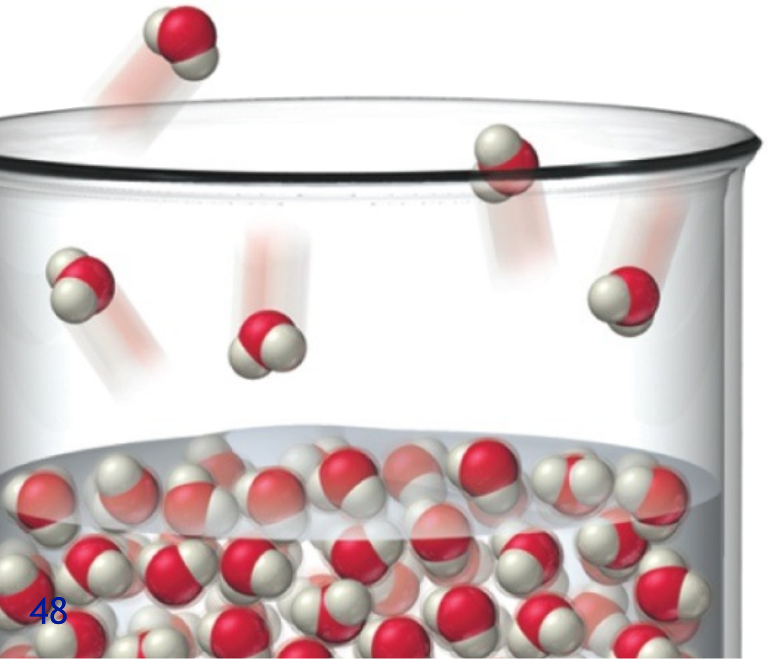


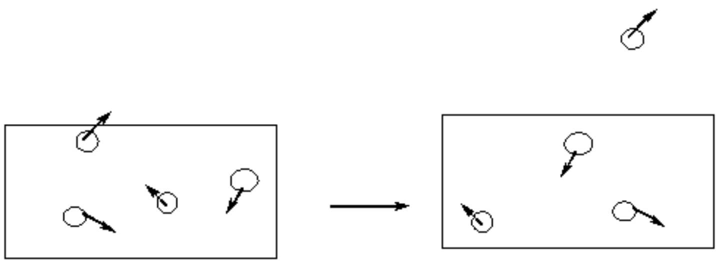


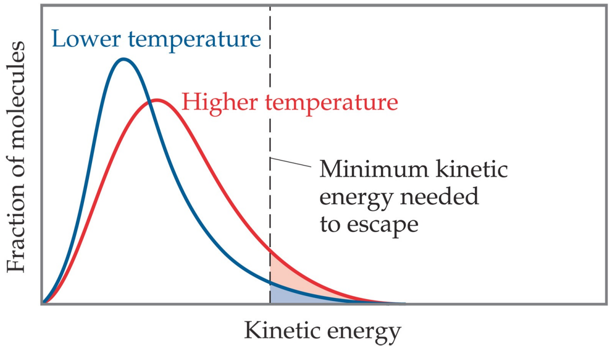


Heat of Vaporization

- Evaporation is an endothermic process.
- Thermal energy is motion of particles.
  - Temperature is a measure of the average speed of particles in a system.
  - As the faster moving (hotter) particles escape the system
    - … the average speed (temperature) of the remaining particles decreases.
- When particles evaporate from a container, the net result is lower the average energy of that container. It get's colder.
- We can measure that change in energy.
- The energy lost with vaporizing one mole of a substance is the heat of vaporization (ΔH_{vap}).



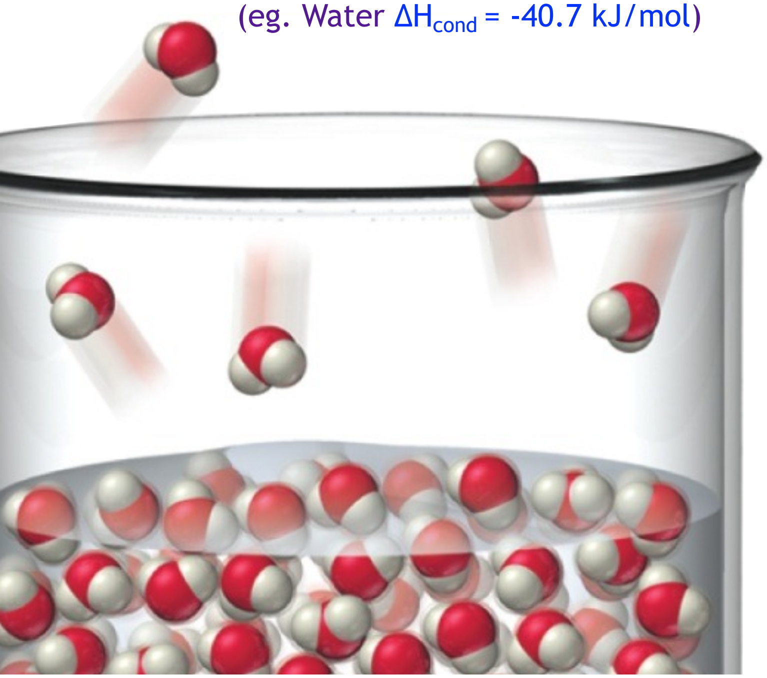




Heat of Vaporization

- The enthalpy required to convert one mole of a substance from liquid to vapor is it's heat of vaporization (ΔH_{vap}).
    - It's often provided in kJ/gram as well — watch the units!
- Vaporization is endothermic (absorbs heat).
- Heat of vaporization is a unique physical property.
    - It's constant for that substance
        - at a given pressure (almost always, assume 1 atm)
    - Example: Water ΔH_{vap} = 40.7 kJ/mol
- Heat of Vaporization is a conversion factor.
- Condensation is the heat released in the reverse process.
- The heat of condensation is the same as the heat of vaporization but exothermic.

  (eg. Water ΔH_{cond} = -40.7 kJ/mol)



How much heat is required to vaporize 2.3 moles of water?

$$q = \Delta H_{vap} \cdot \text{moles}$$
$$= 40.7\ \text{kJ/mol} \cdot 2.3\ \text{mol}$$
$$= 94\ \text{kJ}$$

How much heat is released when 6.7 grams of water is condensed?

$$6.7\ \text{grams} \cdot \frac{1\ \text{mole}}{18.02\ \text{g}} = 0.37\ \text{moles}$$

$$q = \Delta H_{cond} \cdot \text{moles}$$
$$= -40.7\ \text{kJ/mol} \cdot 0.37\ \text{moles}$$
$$= -15\ \text{kJ}$$

15 kJ released!

Ch11

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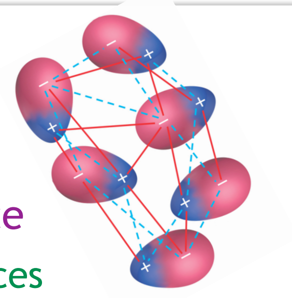

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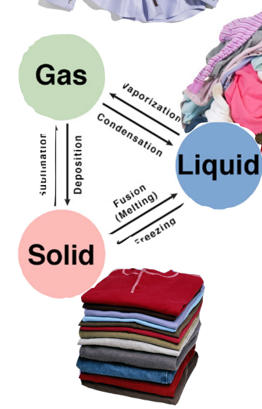
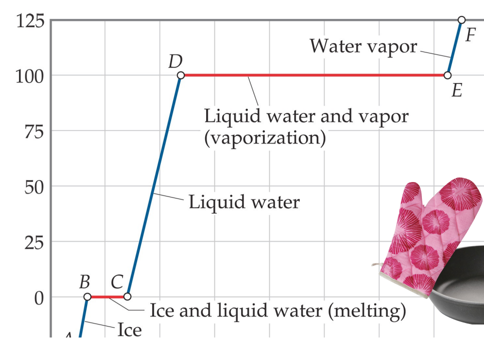
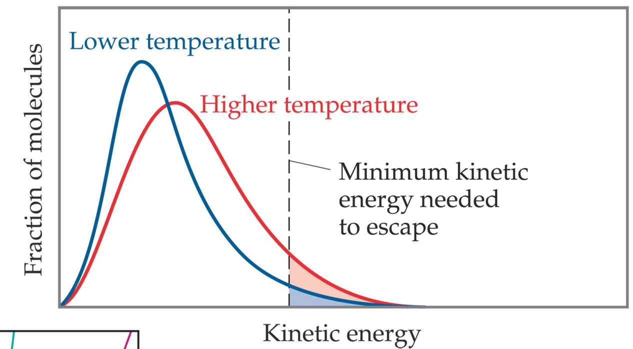
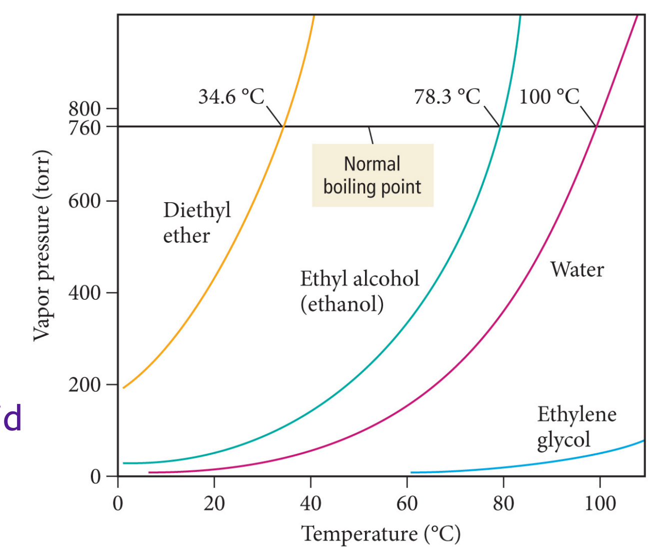




Boiling Point

- Boiling point is the temperature at which the Vapor pressure of the liquid equals the pressure of the world around it.
- It's a threshold.
- It's when a substance has absorbed as much heat as it possibly can in that state.
- It's the point where the heat absorbed is equal to the intermolecular forces holding that substance together.
- Every joule of heat that goes into that substance after it's reached it's boiling point defeats the intermolecular forces of some portion of that substance.
- It vaporizes that portion.
- It explodes it.
- And any portion still a liquid, stays at the boiling point.
- So water boils at 100˚C.
- As you add more heat, you vaporize more of that water.
- But the temperature stays a dependable, unwavering 100˚C until there is absolutely no liquid water left.





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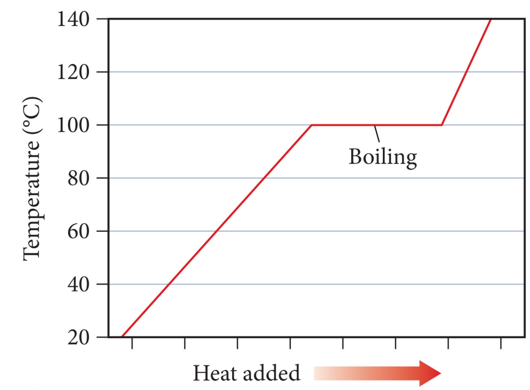

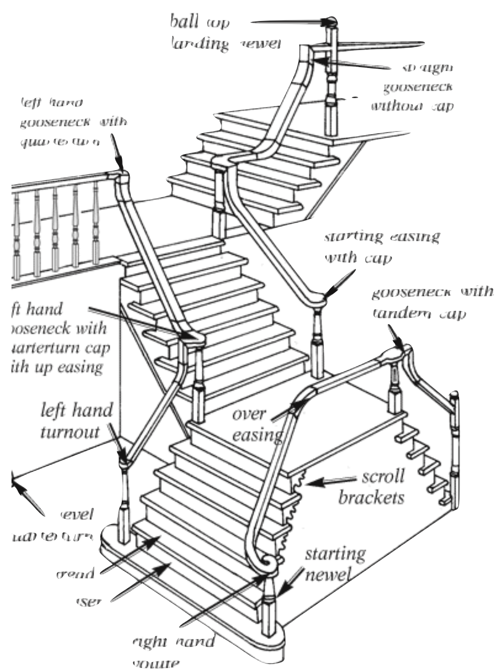

Ch11

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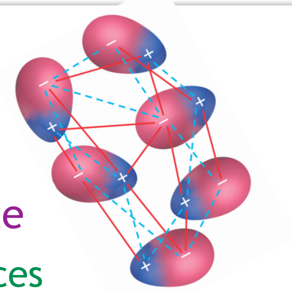

- Properties of Condensed Matter
  - IMF Controls Many Properties
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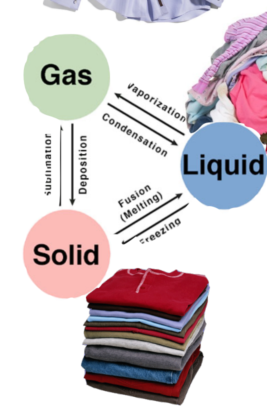
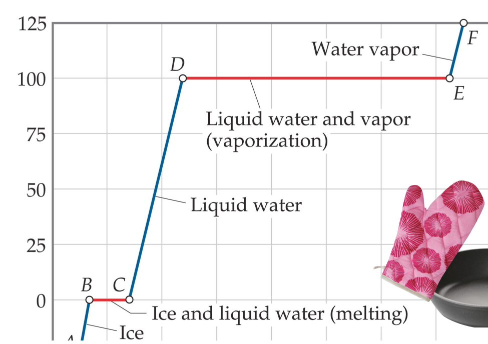
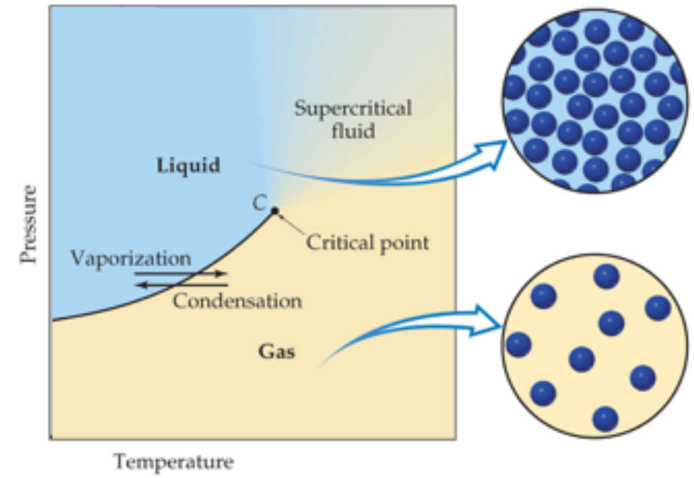
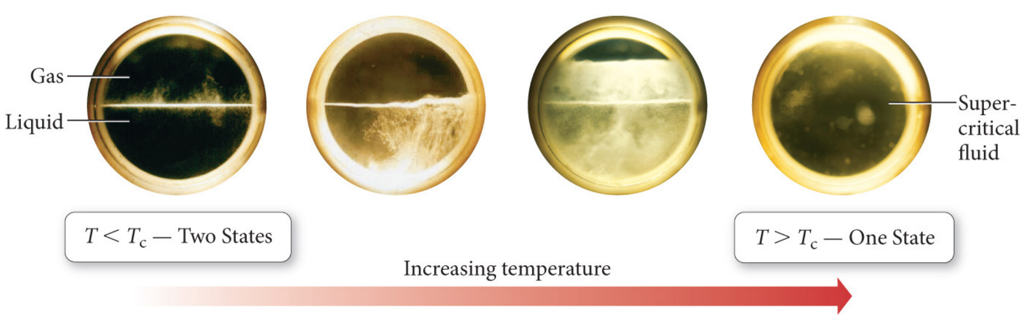




Critical Temperature

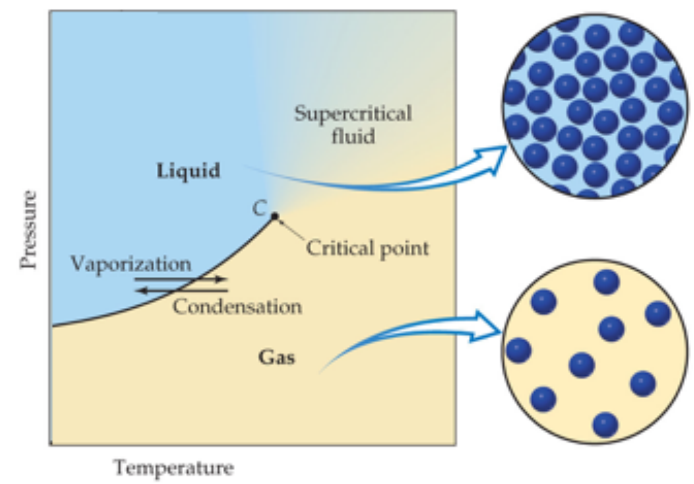
- At a high enough temperature, liquid and gas phases end.
- Regardless of pressure, molecules have so much energy they ignore intermolecular forces.
- The temperature at which substances cannot form liquids is the critical temperature of that substance.

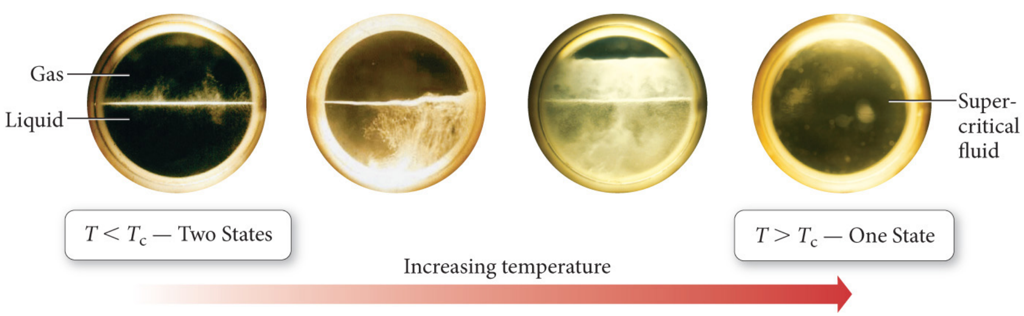




Critical Temperature

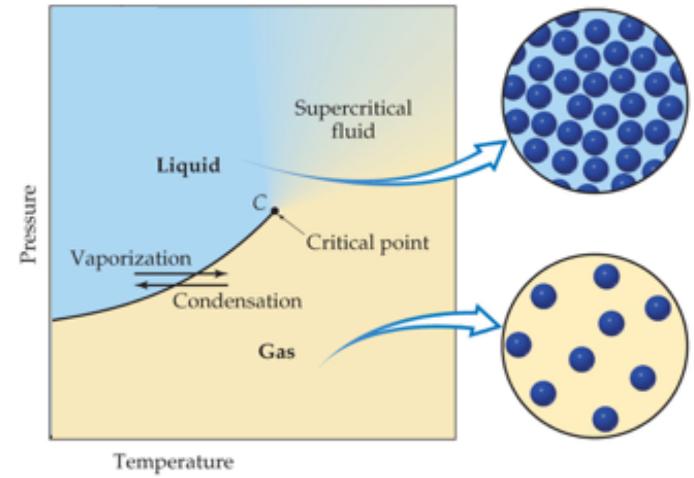
- The pressure at which this transition occurs is the critical pressure of the substance.
- The point defined by the critical temperature and critical pressure of a substance is the critical point.
- Substances above their critical point are said to be neither gas nor liquid, but are described as a supercritical fluid.

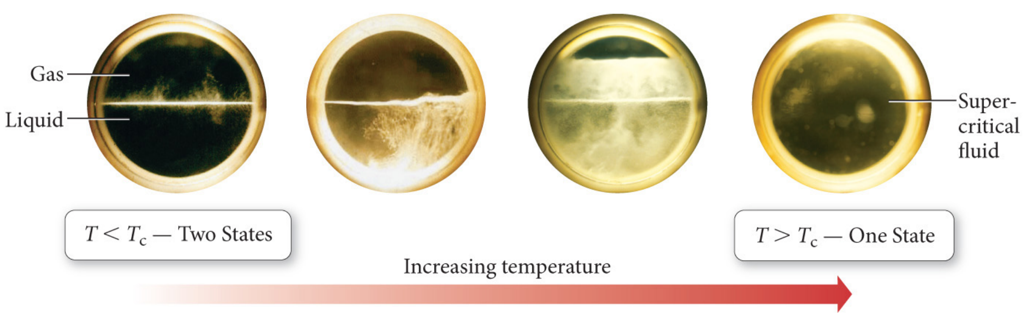




Critical Temperature

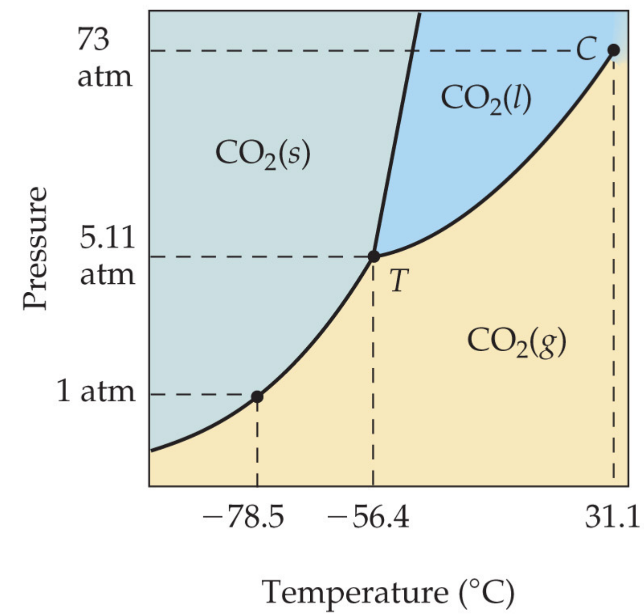
- As a liquid is heated in a sealed container, more vapor collects, causing the pressure inside the container to rise, the density of the vapor to increase, and the density of the liquid to decrease.
- At some temperature, the meniscus between the liquid and vapor disappears, and the states commingle to form a supercritical fluid.
- Supercritical fluids have properties of both gas and liquid states.

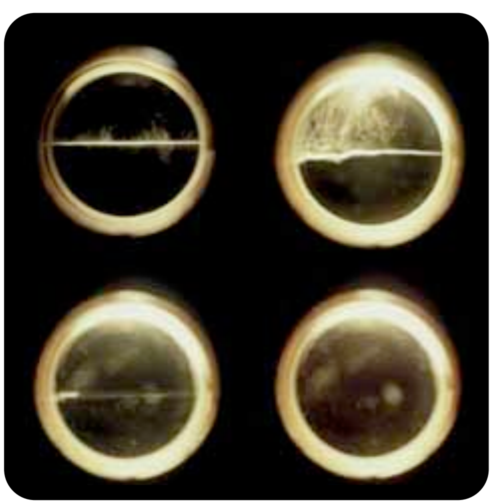


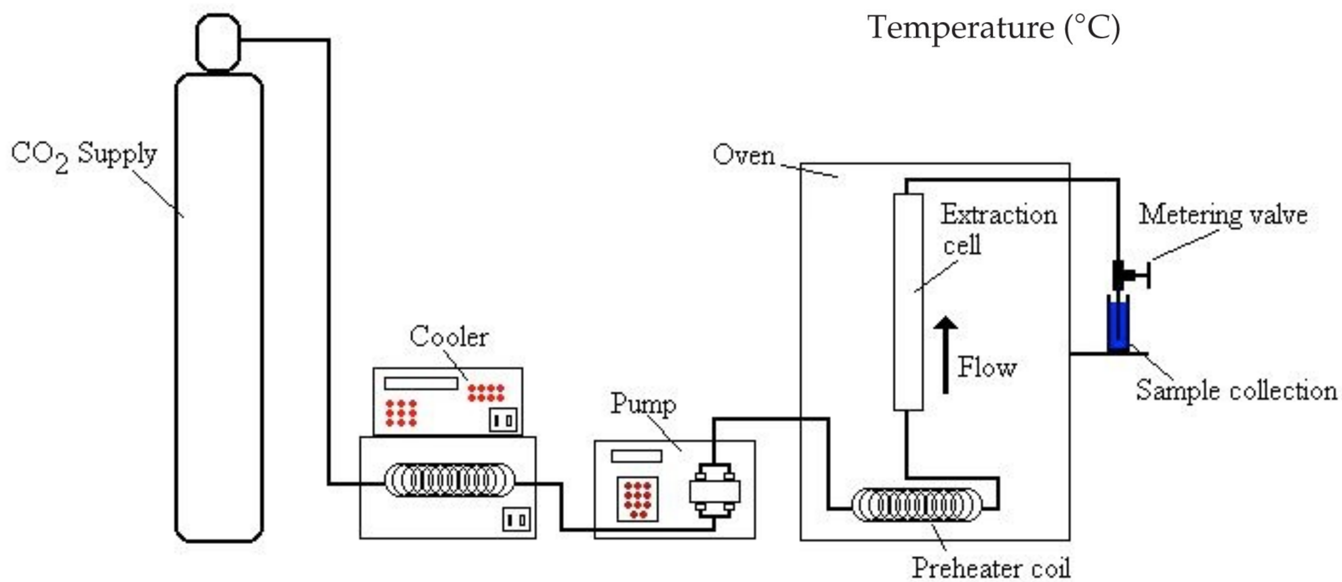


Super Critical Fluids

- The low critical temperature and critical pressure for CO_2 make supercritical CO_2 a unique solvent for extracting nonpolar substances (like caffeine).
- Supercritical fluids behave like both liquids and solvents.
- They mix themselves at gas speeds, they solvate materials like a liquid.
- Placing raw material in supercritical CO_2 causes caffeine to be extracted into the liquid.
- Decanting the CO_2 solutions and releasing the pressure causes the solvent to distill itself off leaving the pure caffeine behind.







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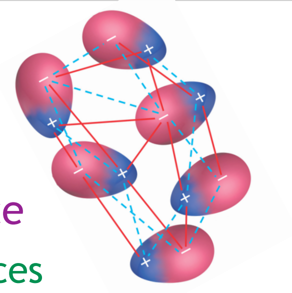

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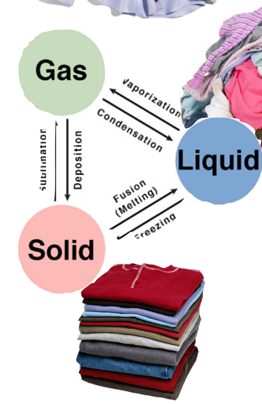
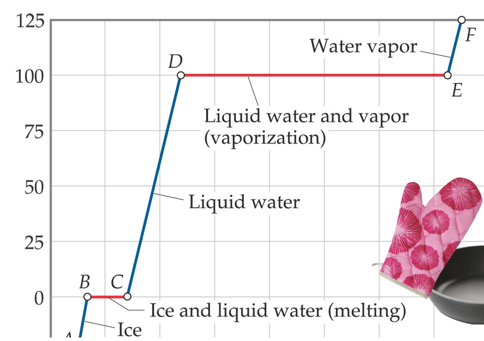
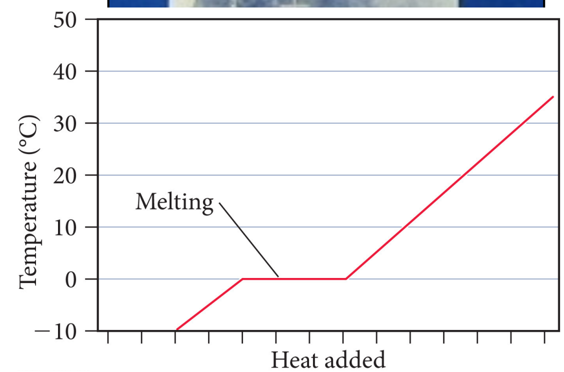
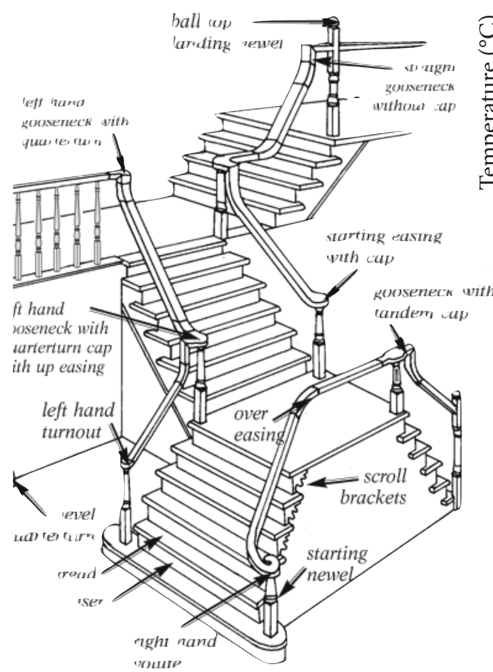




Melting

- Melting point is a second threshold.
- Melting point is where the intermolecular forces that lock the substance into a rigid solid are overcome by the heat added.
- It's when the substance has absorbed all the heat it can possibly hold in that rigid solid structure.
- Every joule of heat added after that, takes some portion of solid to a liquid – at the same temperature.
- Ice baths:
  - An ice bath contains water at 0˚C and ice at 0˚C.
    - Add more heat, more of it becomes water – but the temperature doesn't change.
    - Remove heat, more of it becomes ice – but the temperature doesn't change.
  - That's why we use ice baths – to insure a constant, reliable temperature.

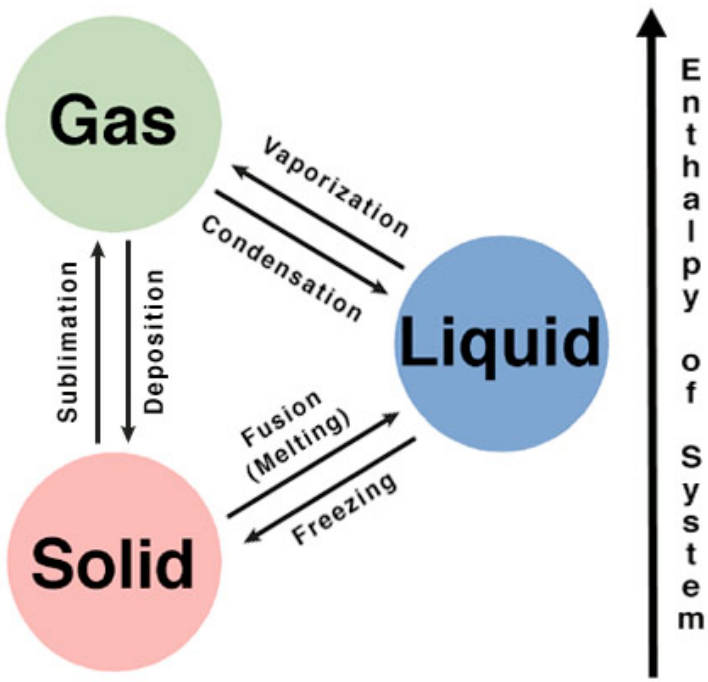




State Changes

- There are three states of matter and six processes by which we change state.
- Three are endothermic processes, three are complementary exothermic processes.

Note: The process of melting is described with Heat of Fusion. (It's an old convention)



Endothermic

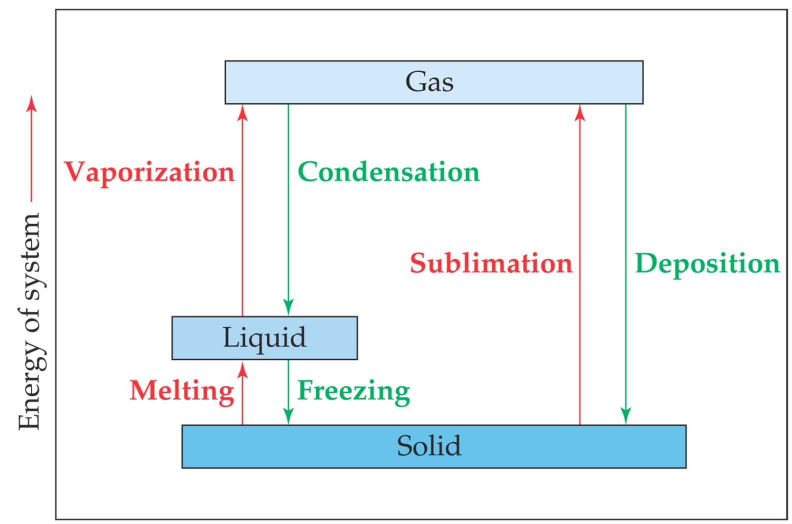
- Fusion
- Vaporization
- Sublimation
- Condensation
- Freezing
- Deposition

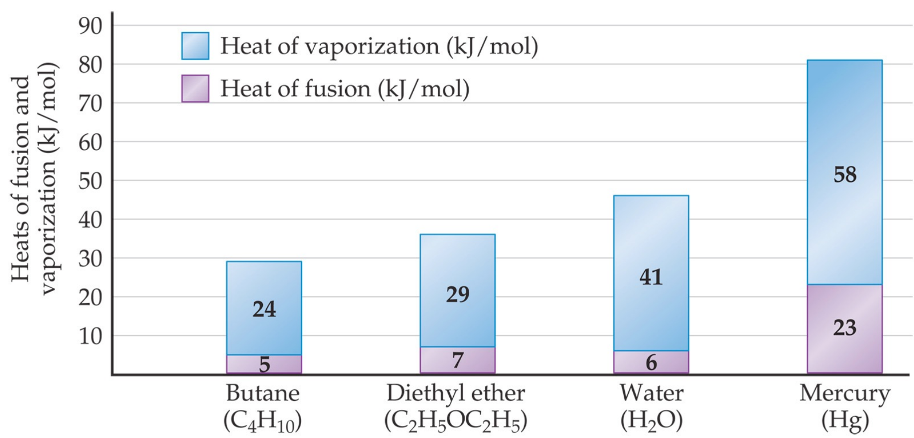
Exothermic

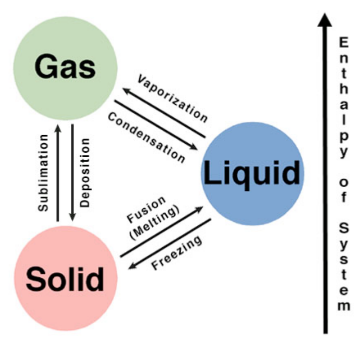
State Changes

- There are three states of matter and six processes by which we change state.
- Three are endothermic processes, three are <u>complementary</u> exothermic processes.
- Because of Hess's Law, Heat of Sublimation is equal to the sum of the heat of Vaporization and Fusion.

- Fusion ΔH_{fus} e.g. 20.2 kJ/mol
- Vaporization ΔH_{vap} e.g. 58.2 kJ/mol
- Sublimation ΔH_{sub} e.g. 78.4 kJ/mol
- Freezing $\Delta H_{frz} = -\Delta H_{fus}$ e.g. -20.2 kJ/mol
- Condensation $\Delta H_{con} = -\Delta H_{vap}$ e.g. -58.2 kJ/mol
- Deposition $\Delta H_{dep} = -\Delta H_{sub}$ e.g. -78.4 kJ/mol







Ch11

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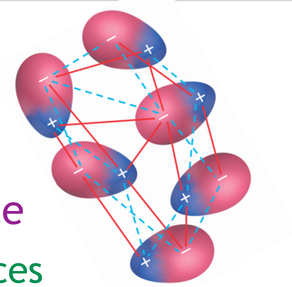

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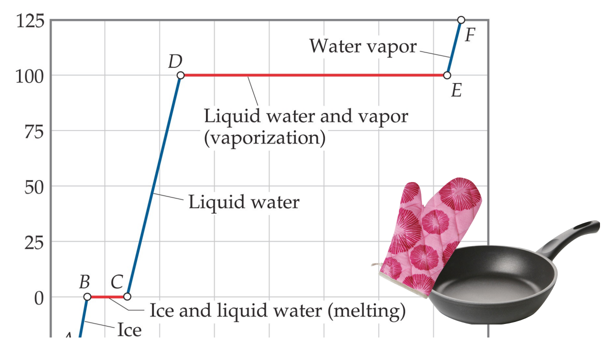
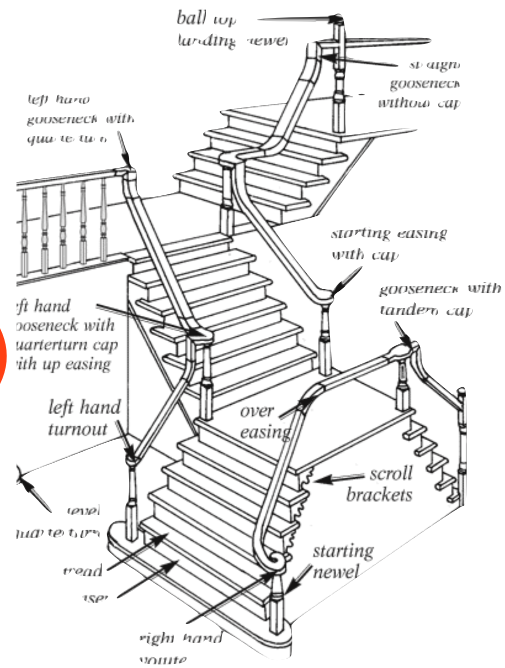


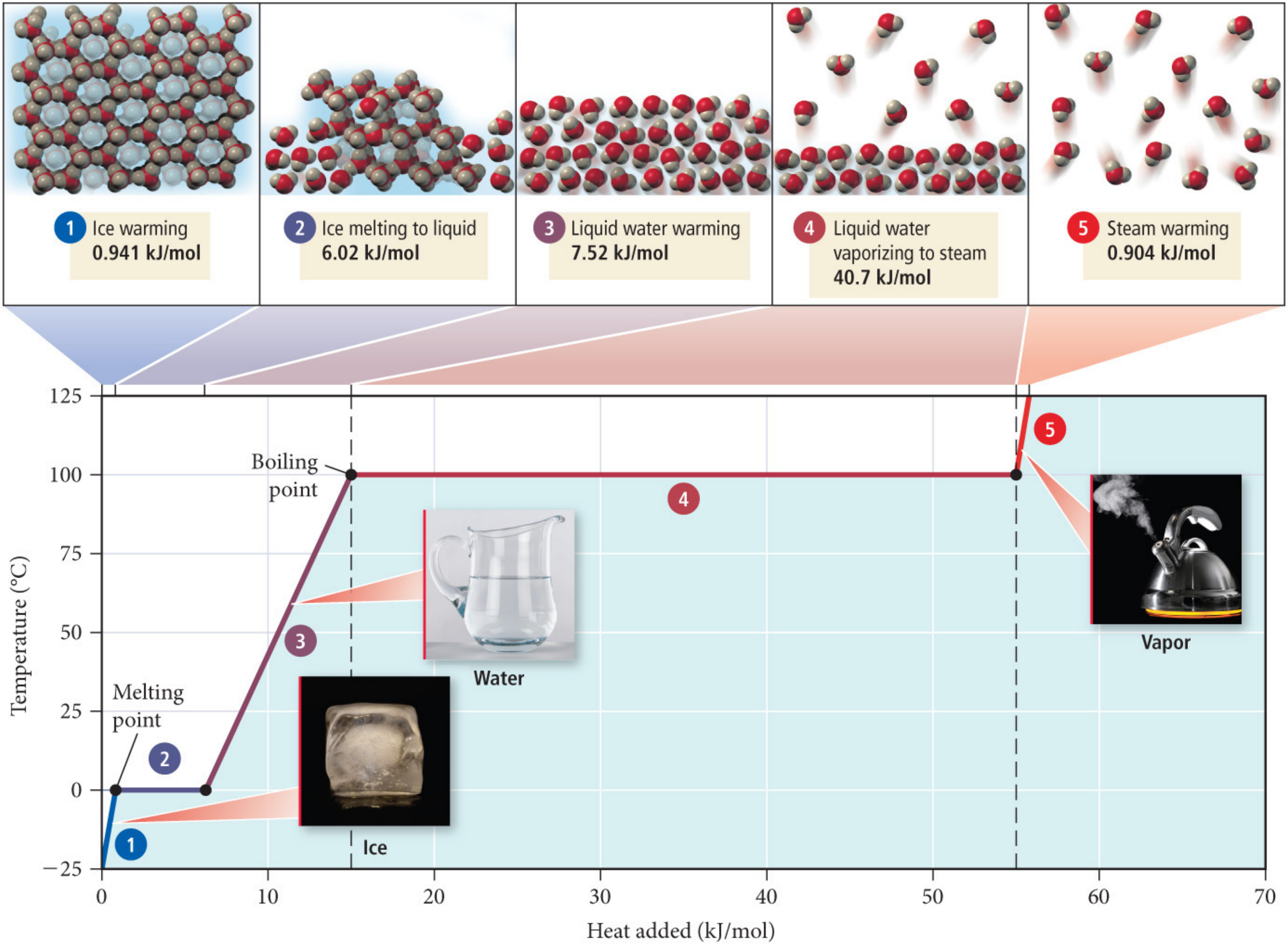
Heat Effects Matter

- Heat (q) has two principle effects on matter.
  - It can cause a change in temperature.
  - It can cause matter to change state.
- Which effect occurs, *what the heat does*, depends on where you are relative to threshold temperatures.
    - Melting Point
    - Boiling Point

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚- 100˚C	Liquid Changes Temperature
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature



Heating Water and Boiling Water are very different processes!


1 Ice warming
0.941 kJ/mol
2 Ice melting to liquid
6.02 kJ/mol
3 Liquid water warming
7.52 kJ/mol
4 Liquid water vaporizing to steam
40.7 kJ/mol
5 Steam warming
0.904 kJ/mol
Temperature (°C)
125
100
75
50
25
0
−25
Boiling point
Melting point
Water
Ice
Vapor
0
10
20
30
40
50
60
70
Heat added (kJ/mol)

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  - It can cause matter to change state.
- Which effect occurs, *what the heat does*, depends on where you are relative to threshold temperatures.
    - Melting Point
    - Boiling Point
- Heat Capacity allows you calculate the heat required to change temperature.
- Heat of Fusion allows you to calculate the heat required to change state.
- Both equations require you to consider the quantity of the substance.
  - The quantity can represented as
    - moles (n) or mass (m).
- In either case you need the appropriate conversion factor.
  - C_s (specific heat for mass)
  - C_m (molar heat capacity for moles)
  - ΔH_s (specific heat of fusion for mass)
  - ΔH_m (molar heat of fusion for moles)

Temperature	Effect on Water
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At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

Heat Capacity

$$q = C\,\Delta T$$

$$C = C_s \cdot m$$
$$C = C_m \cdot n$$

Heat of Fusion

$$q = \Delta H$$

$$\Delta H = \Delta H_s \cdot m$$
$$\Delta H = \Delta H_m \cdot n$$

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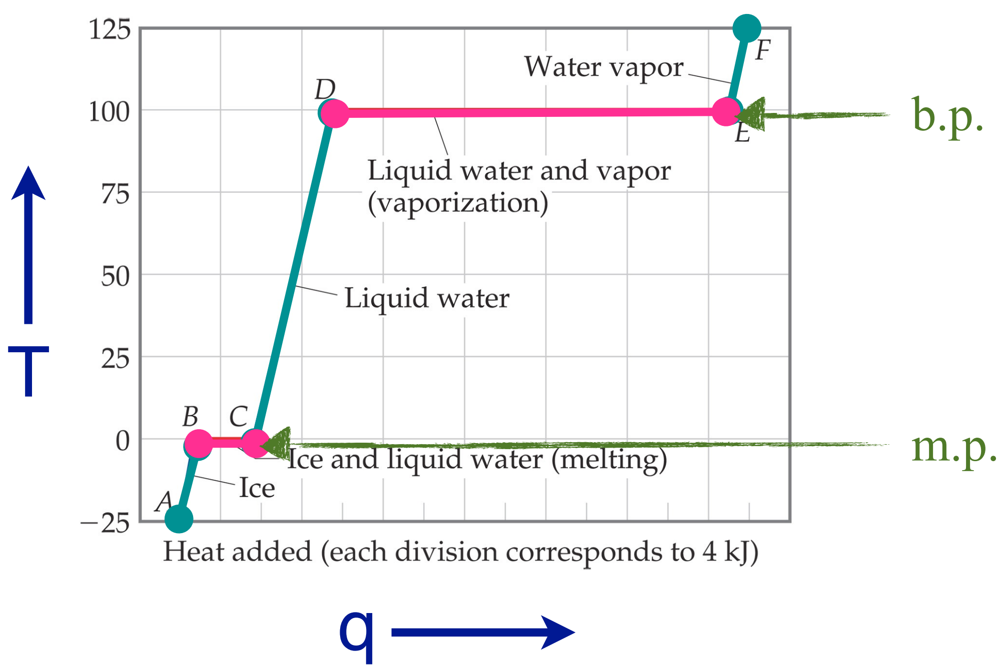
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature



Heat Capacity

$$q = C\,\Delta T$$

$$C = C_s \cdot m$$
$$C = C_m \cdot n$$

Heat of Fusion

$$q = \Delta H$$

$$\Delta H = \Delta H_s \cdot m$$
$$\Delta H = \Delta H_m \cdot n$$

Problem:

What is the heat in kilojoules required to heat 25.0 grams of ice from -25.0 ˚C to -5.0 ˚C?

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚- 100˚C	Liquid Changes Temperature
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

heat of fusion of water = 334 J/g
heat of vaporization of water = 2257 J/g
specific heat of ice = 2.09 J/g·°C
specific heat of water = 4.18 J/g·°C
specific heat of water vapor = 1.99 J/g·°C

Specific Heat

$$q = C_s \bullet m \bullet \Delta T \ ; \ J \rightarrow kJ$$

m = 25.0 grams

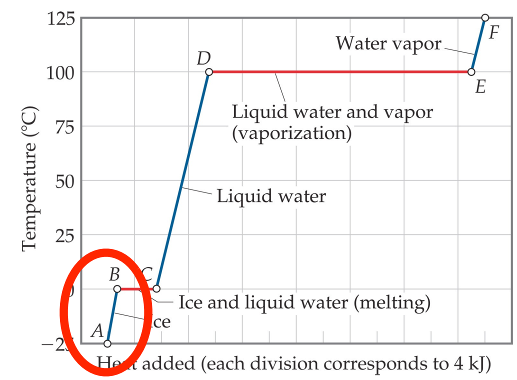
q = heat required

$C_{s\ ice}$= 2.09 J/g˚C

T_F = -25.0 ˚C
T_I = -5.0 ˚C

ΔT = 20.0 ˚C

(ΔT is positive because you should be doing this in Kelvin)



$$q = 2.09\ \frac{J}{g^{\circ}C} \times 25.0\ g \times 20^{\circ}C = 1045\ J$$

$$q = 1045\ J \times \frac{1kJ}{1000J} = \boxed{1.05\ kJ}$$

Problem:

What is the heat in kilojoules required to melt 25.0 grams of ice?

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚ - 100˚C	Liquid Changes Temperature
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

heat of fusion of water = 334 J/g
heat of vaporization of water = 2257 J/g
specific heat of ice = 2.09 J/g·°C
specific heat of water = 4.18 J/g·°C
specific heat of water vapor = 1.99 J/g·°C

Heat of Fusion

$$q = \Delta H_s \bullet m \text{ ; J} \rightarrow \text{kJ}$$

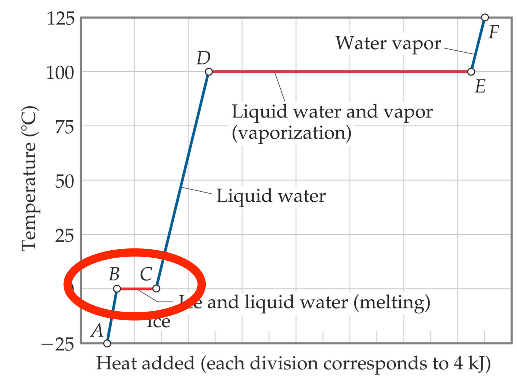
m = 25.0 grams

q = heat required

ΔH_{fus} = 334 J/g

$$q = 334 \frac{J}{g} \times 25.0 \text{ g}$$

$$q = 8350 \text{ J} \times \frac{1 \text{kJ}}{1000 \text{J}} = 8.35 \text{kJ}$$



Problem:

What is the heat in kilojoules required to heat 25.0 grams of water from 0.0˚C to 73.8 ˚C?

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚- 100˚C	Liquid Changes Temperature
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

heat of fusion of water = 334 J/g
heat of vaporization of water = 2257 J/g
specific heat of ice = 2.09 J/g·°C
specific heat of water = 4.18 J/g·°C
specific heat of water vapor = 1.99 J/g·°C

Specific Heat

$$q = C_s \bullet m \bullet \Delta T \; ; \; J \rightarrow kJ$$

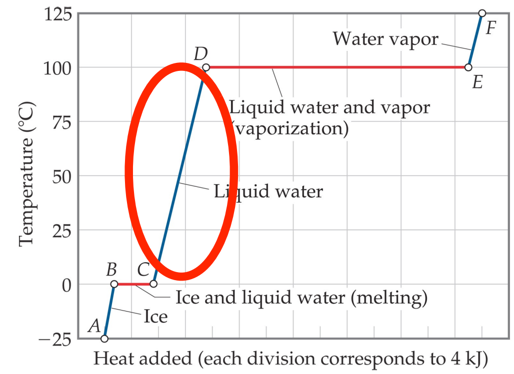
m = 25.0 grams

q = heat required

$C_{s\ water} = 4.18\ J/g˚C$

$T_F = 73.8\ ˚C$
$T_I = 0.0\ ˚C$

$\Delta T = 73.8\ ˚C$



$$q = 4.18\ \frac{J}{g°C} \times 25.0\ g \times 73.8°C = 7712.1\ J$$

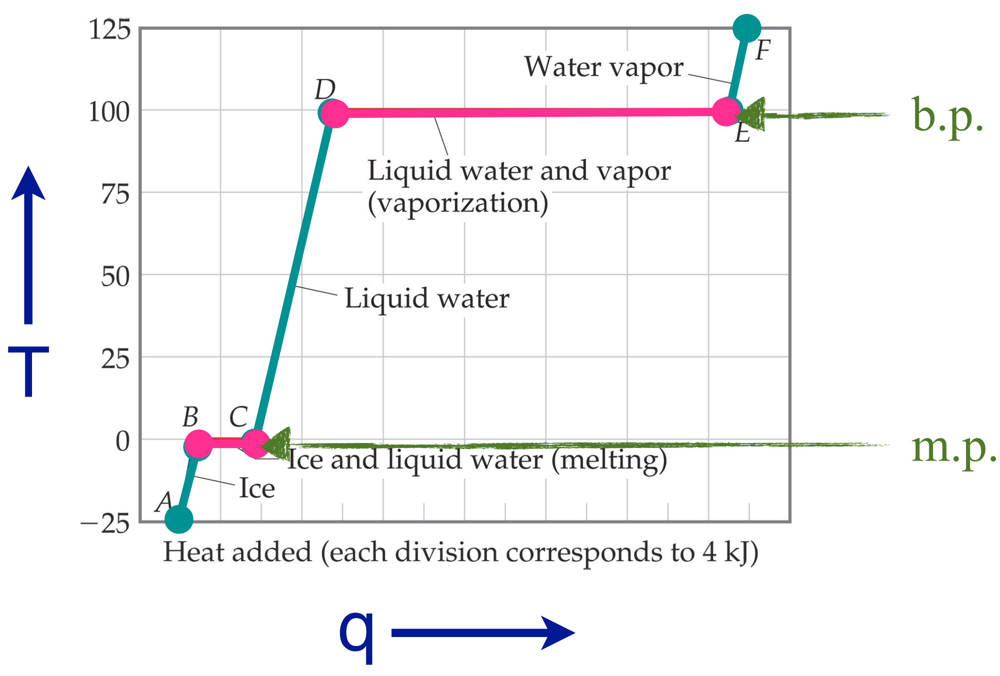
$$q = 7712.1\ J \times \frac{1kJ}{1000J} = \boxed{7.71\ kJ}$$

Heat Effects Matter

- Heat (q) has two principle effects on matter.
  - It can cause a change in temperature.
  - It can cause matter to change state.
- Which effect occurs, *what the heat does*, depends on where you are relative to threshold temperatures.
- Heat is a state function. It's additive.

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚ - 100˚C	Liquid Changes Temperature
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

$$q_{B\text{-}D} = q_{B\text{-}C} + q_{C\text{-}D}$$



Heat Capacity

$$q = C\,\Delta T$$

$$C = C_s \cdot m$$
$$C = C_m \cdot n$$

Heat of Fusion

$$q = \Delta H$$

$$\Delta H = \Delta H_s \cdot m$$
$$\Delta H = \Delta H_m \cdot n$$

Problem:

What is the heat in kilojoules required to heat 25.0 grams of ice from 0.0˚C to 73.8 ˚C?

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
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specific heat of water = 4.18 J/g·°C
specific heat of water vapor = 1.99 J/g·°C

Heat of Fusion $q_1 = \Delta H_s \cdot m$

Specific Heat $q_2 = C_s \cdot m \cdot \Delta T$

$q = q_1 + q_2$; J → kJ

m = 25.0 grams

q = heat required

ΔH_{fus} = 334 J/g

$C_{s\ water}$ = 4.18 J/g˚C

T_F = 73.8 ˚C

T_I = 0.0 ˚C

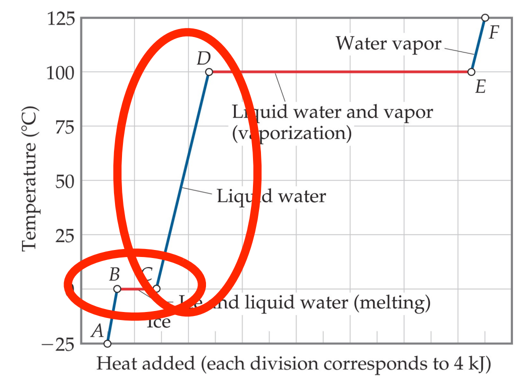
ΔT = 73.8 ˚C

$$q_1 = 334 \frac{J}{g} \times 25.0\ g = 8350\ J$$

$$q_2 = 4.18 \frac{J}{g^\circ C} \times 25.0\ g \times 73.8^\circ C = 7712.1\ J$$

$$q = q_1 + q_2 = 8350\ J + 7712.1\ J = 1{,}6062.1\ J$$

$$q = 1{,}6062.1\ J \times \frac{1kJ}{1000J} = \boxed{16.06\ kJ}$$



$$\begin{array}{r} 8350 \\ 7710 \\ \hline 16060 \end{array}$$

A Steam Engine

Water at 45.5˚ is heated to convert it entirely to steam at exactly 100˚C. How much heat is needed to convert 23.2 grams?

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚ - 100˚C	Liquid Changes Temperature
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

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heat of vaporization of water = 2257 J/g
specific heat of ice = 2.09 J/g·°C
specific heat of water = 4.18 J/g·°C
specific heat of water vapor = 1.99 J/g·°C

45.5–100° – Liquid water Heats

$T_f = 100.0$
$T_I = 45.5\,°C$
$\Delta T\ 54.5\,°C$

$m = 23.2\,g$

$C_s = 4.18\ J/g°C$

$q_1 = C_s \cdot m \cdot \Delta T$
$= 4.18\ J/g°C \cdot 23.2\,g \cdot 54.5\,°C$
$= 5{,}285.192\ J$
$= 5{,}290\ J$

Total Heat = $q_1 + q_2$

$= 57{,}700\ J$

100° – Liquid water changes to Steam

$m = 23.2\,g$

$\Delta H_{VAP} = 2257\ J/g$

$q_2 = \Delta H_{VAP} \cdot mass$
$= 2257\ J/g \cdot 23.2\,g$
$= 52{,}362.40\ J$
$= 52{,}400\ J$

5290
+ 52400
57690

Melting Ice

A 35.0 g block of metal at 80.0˚C is added to a mixture of 100.0 g water and 15.0 g of ice in an isolated container. All the ice melted and the temperature in the container rose to 10.0 ˚C.

How much heat came out of the metal?

Temperature	Effect on Water
Above 100˚C	Gas Changes Temperature
At exactly 100˚C	Liquid Converts to Gas
Between 0˚ - 100˚C	Liquid Changes Temperature
At exactly 0˚C	Solid Converts to Liquid
Below 0˚C	Solid Changes Temperature

heat of fusion of water = 334 J/g
heat of vaporization of water = 2257 J/g
specific heat of ice = 2.09 J/g·°C
specific heat of water = 4.18 J/g·°C
specific heat of water vapor = 1.99 J/g·°C

Heat out of metal = Heat into ice/water
Ice Water is 0°C (exactly)

Melts 15.0 g ice at 0°C

mass = 15.0 g
ΔH_{fus} = 334 J/g

$q_1 = \Delta H_{fus} \cdot mass$
$= 334\ J/g \cdot 15.0\ g$
$= 5{,}010.00$
$= 5{,}010\ J$

5,010
4,810
9,820

$q_r = q_1 + q_2$
$= 9{,}820\ J$

$= 9.82\ kJ$

$T_F = 10.0°C$
$T_I = 0°C$ (exact)
$\Delta T = 10.0°C$

$C_s = 4.18\ J/g°C$

mass = 100.0 + 15.0 g
= 115.0 g

HEATING WATER 0°C → 10.0°C

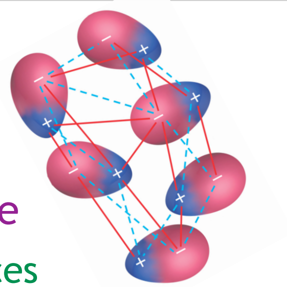
$q_2 = C_s \cdot m \cdot \Delta T$
$= 4.18\ J/g°C \cdot 115.0\ g \cdot 10.0°C$
$= 4{,}810\ J$

Ch11

Condensed Matter

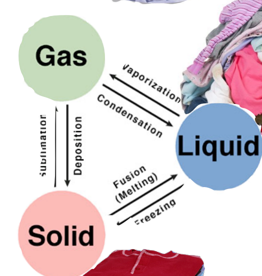
inTRA — Within a molecule
inTER — Between molecules

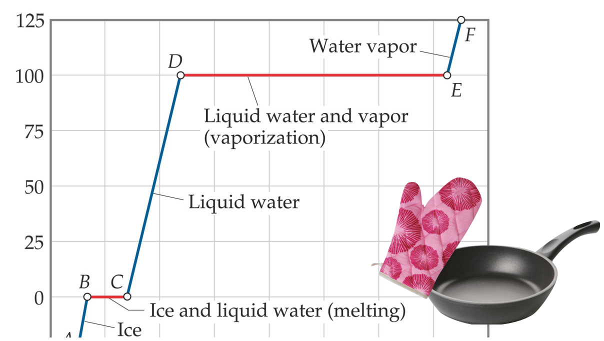
- Solids & Liquids
  - Sticky Molecules
  - Intermolecular Force
    - Dipole Dipole Forces
      - Scales with polarity
    - London Forces
      - Scales with molecular size & shape
    - Hydrogen Bonding (O, N, F)
- Properties of Condensed Matter
  - IMF Controls Many Properties
  - When IMF & Energy Balance
    - Fluidity, Viscosity
  - Cohesion & Adhesion
    - Surface Tension & Droplet Shape
    - Capillary Action
    - Meniscus Shape





- Characteristic Temperatures
  - Vaporization/Condensation
    - at the Boiling Point [ΔH_{vap}]
  - Critical Temperature
  - Sublimation/Fusion
    - at the Melting Point [ΔH_{fus}]
- Heat & Matter
  - Connecting the pieces…
    - T < mp (heating solids)
    - T = mp (melting solids)
    - bp > T > mp (heating liquids)
    - T = bp (boiling liquids)
    - T > bp (heating gases)

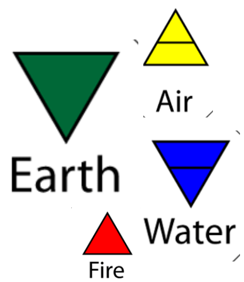



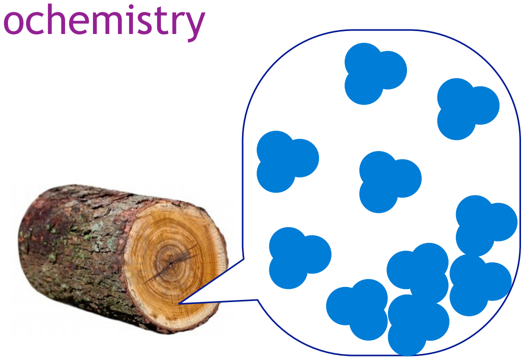
Questions?

Chem 210 is Complete

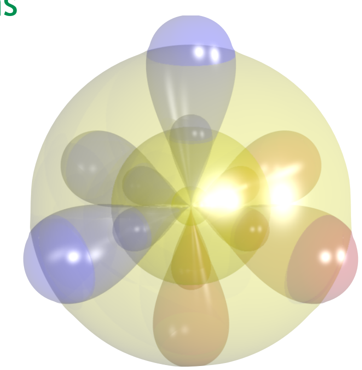
General Chemistry part 1

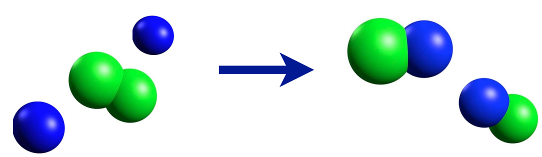
- Ch 1
  - Science, Matter, & Chemistry
- Ch 2
  - Atoms & Elements
- Ch 3
  - Molecules & Compounds
- Ch 4
  - Solutions
- Ch 5
  - The Gas State of Matter
- Ch 6
  - Thermochemistry





- Ch 7
  - Quantum Mechanics
    - The structure of the Atom
- Ch 8
  - Electronic Configurations
    - Properties of Atoms
- Ch 9
  - Chemical Bonds
- Ch 10
  - Molecular Shape
- Ch 11
  - Intermolecular Forces
    - The Condensed States of Matter





The End