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Intermolecular Forces

Love & Hate in the Molecular Realm

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If I put 2 molecules into a sealed flask, what could happen?

1. They ignore each other.
2. They LOVE each other – they're attracted to each other
3. They HATE each other – they repel each other

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REMEMBER:
MOLECULES MOVE!

(except at 0 K)

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If they LOVE each other, what would that look like?

Initially Later

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If they HATE each other, what would that look like?

Initially Later

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If they IGNORE each other, what would that look like?

Initially Later

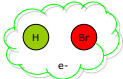
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What determines LOVE or HATE?

The structure of the molecule.

What is the structure of a molecule?



What's in the nuclei?
Protons!

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Molecular structure is all about...

POSITIVE & NEGATIVE CHARGES!

So Love & Hate is all about...

Opposites attract, like repel!

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Types of Intermolecular Forces

1. London Dispersion forces, aka Van der Waal's forces, aka Instantaneous dipole-induced dipole forces.
2. Dipole-Dipole interactions
3. Hydrogen bonding – particularly strong case of dipole-dipole interaction
4. Ionic forces
5. Mixed forces

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London Dispersion forces, aka van der Waal's forces, aka Instantaneous dipole-induced dipole forces.

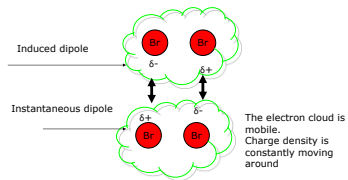
This is NOT the strongest, but it is the primary intermolecular force.

All atoms or molecules with electrons have Van der Waal's forces – so ALL atoms or molecules have Van der Waal's forces

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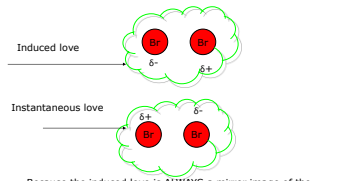
Instantaneous dipole-induced dipole forces



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How Great is THAT!?!?!?



Because the induced love is ALWAYS a mirror image of the instantaneous love, dispersion forces are ALWAYS attractive

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Dispersion Forces are ALWAYS ATTRACTIVE

All molecules like each other, at least a little bit. So all molecules stick together, at least a little bit.

If they didn't...
...the universe would be a much more chaotic place!
Occasional repulsion would have things flying apart all over the place!

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Van der Waal's forces

Van der Waal's forces get stronger as the temporary dipole gets stronger.

The temporary dipole is caused by electron mobility, so the more electrons the stronger the Van der Waal's forces.

electrons increases as # protons, so the heavier the molecule the stronger the Van der Waal's forces.

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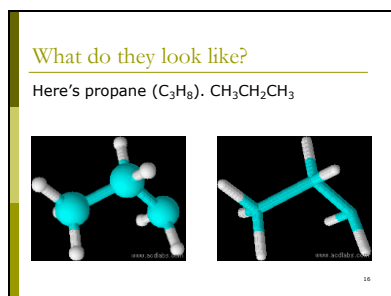
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Alkanes

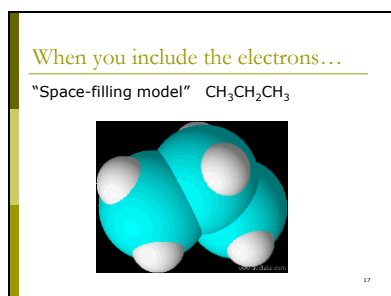
Methane - CH_4
Ethane - CH_3CH_3
Propane - $\text{CH}_3\text{CH}_2\text{CH}_3$
Butane - $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$
Pentane - $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$
Hexane - $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$
Heptane - $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$
Octane - $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$

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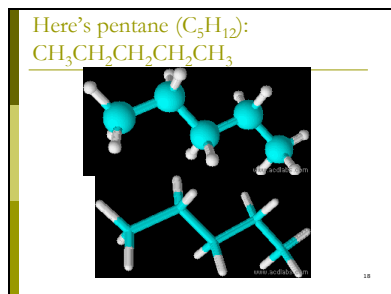
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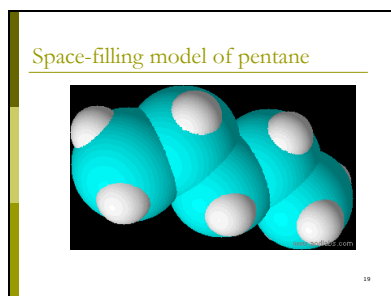
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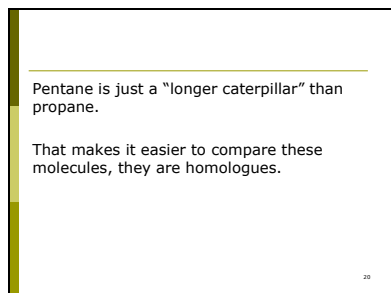
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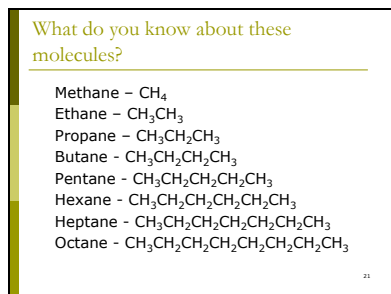
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What do you know about these molecules?

Methane – gas at standard T & P
Ethane – gas at standard T & P
Propane – gas at standard T & P – Liquid under slight pressure
Butane – gas at standard T & P – Liquid under slight pressure
Pentane – Liquid
Hexane – Liquid
Heptane – Liquid
Octane – Liquid

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Solids, Liquids, and Gases

What is the difference between a solid, a liquid, and a gas microscopically?

How tightly stuck together the molecules are!!!

Solids are stuck together more than liquids that are stuck together more than gases

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Sticking together is a function of TWO things:

How much you like each other...and how much you are trying to get away from each other.

In the context of molecules, this is a question of intermolecular forces vs. kinetic energy.

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ALL MOLECULES
MOVE!

(except at 0 K)

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Motion = Kinetic Energy =
Temperature

Kinetic energy is energy of motion.

Temperature is a measure of the "mean kinetic energy of molecules".

Temperature reflects your desire to escape...

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Solids, Liquids, and Gases & Heat

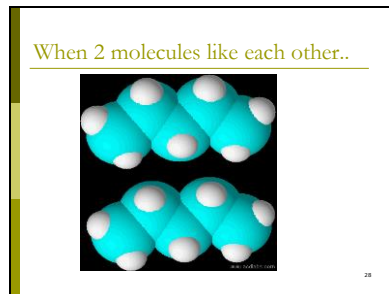
What happens when you heat up a solid?

Eventually it melts – why?

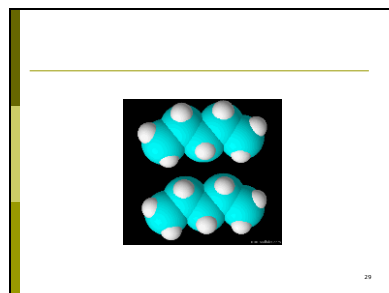
Adding heat adds energy to the molecules, when they have enough energy they can escape their attraction to their neighbors!

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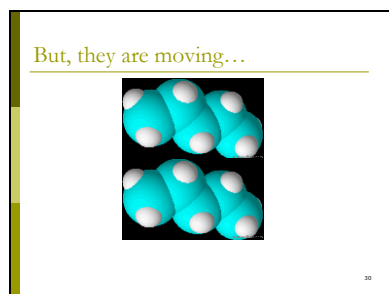
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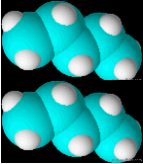


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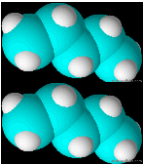
If they are HOT HOT HOT!



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If they are very COLLLLDDD



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Intermolecular forces...

...only depend on distance between the molecules.

I can't change the structure of the molecule.
But the farther apart they are, the smaller the force they feel.

[Think gravity and distance from the center of the earth.]

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I can add heat...

"Hot" and "Cold" are relative...

Phase is a balance between the temperature (kinetic energy) of the molecules that is trying to separate them and the intermolecular forces which are trying to hold them together.

Melting point or boiling point is the kinetic energy where the balance tips.

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Van der Waal's Forces are...

...the first consideration – but not the last!

But ALL the intermolecular forces are about CHARGE! (Opposites attract.)

ALL intermolecular forces are ATTRACTIVE in the end.

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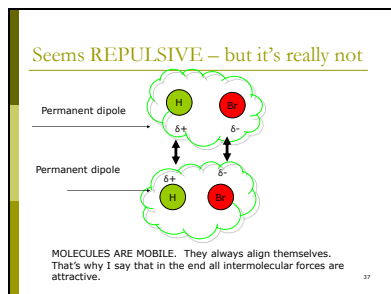
Dipole – Dipole Interactions

Permanent dipole

Permanent dipole

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Dipole – Dipole interactions

A molecule with a permanent dipole is called a "polar molecule".

All polar molecules have Dipole-Dipole interactions in ADDITION TO Van der Waal's forces.

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Dipole – Dipole interactions

Dipole-Dipole interactions are in ADDITION TO Van der Waal's forces.

They are generally weaker and just add on to VDW forces with ONE EXCEPTION.

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Hydrogen Bonding – just a special case of dipole-dipole interactions

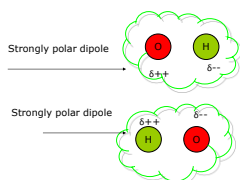
Hydrogen bonding is a dipole-dipole interaction that occurs when hydrogen is bonded to something very electronegative like F, O, or N.

It is just a very strong dipole-dipole interaction because of the very polar nature of the H-F, H-O, or H-N bond.

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Hydrogen Bonding – just a special case of dipole-dipole interactions



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Compare H₂O to H₂S

Which would you expect to have the higher boiling point?

H₂O has a molar mass of 18 g/mol
H₂S has a molar mass of 34 g/mol

Based on Van der Waal's forces alone, H₂S should have the higher boiling point.

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Compare H_2O to H_2S

The boiling point of water is 373 K.

The boiling point of H_2S is 213 K.

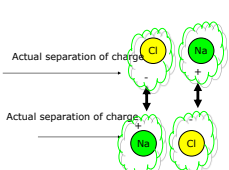
H_2S is a gas at room temperature while water is a liquid!

No FON, no Hydrogen bonding

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Ion – Ion Interactions



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Ion-Ion interactions

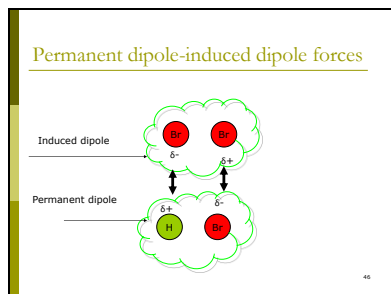
The strongest possible interaction.

The complete charge separation makes it a HUGE dipole-dipole type interaction.

This is why most ionic compounds are solids at room temperature.

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Dipole – Induced Dipole interactions

This is a special case of a Dipole – Dipole interaction where there are 2 different molecules involved and only 1 of them is polar.

Generally weaker than a permanent Dipole-Dipole interaction, it is still IN ADDITION TO Van der Waal's forces.

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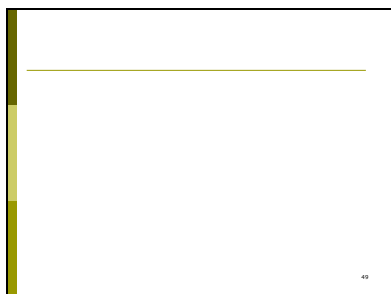
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All the forces...

1. Van der Waal's/Dispersion forces – FIRST consideration. Weakest for single bond BUT it is a more global force. Heavier molecules have bigger VDW forces.
2. Dipole-Dipole forces – add on to VDW forces (with ONE exception - #3). If the molecules have similar mass and shape. The one with a permanent dipole will have a higher boiling point.
3. Hydrogen Bonding TRUMPS VDW
4. Ionic forces TRUMP EVERYTHING

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NaF vs. F₂

What do you know about these 2 molecules?

NaF is an ionic solid

F₂ is a gas at room temp

NaF has a molar mass of 42 g/mol, F₂ has a molar mass of 38 g/mol.

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Ion-ion interactions are the strongest

Based on Van der Waal's forces, you'd expect NaF and F₂ to be similar.

The powerful ionic forces of NaF make it a solid – trumping the Van der Waal's interaction.

NaF melts at 1266 K and boils at 1968 K
F₂ melts at 53 K and boils at 85 K

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HBr vs. Cl₂

What do you know about these 2 molecules?

HBr is a gas at room temp

Cl₂ is a gas at room temp

HBr has a molar mass of 81 g/mol
Cl₂ has a molar mass of 71 g/mol

HBr is polar, Cl₂ is non-polar

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HBr vs Cl₂

So HBr is heavier – more van der Waal's forces

HBr is polar – dipole-dipole forces also

So you would think that HBr has the higher boiling point...and so we go to wikipedia and find...

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HBr boils at 207 K, Cl₂ boils at 239 K

WTWikipedia?!?!?

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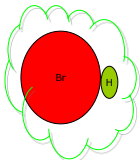
So, why doesn't it?

Geometry is also an issue!

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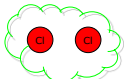
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Geometry



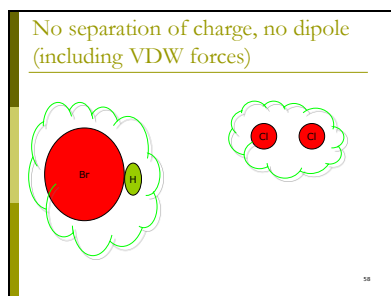
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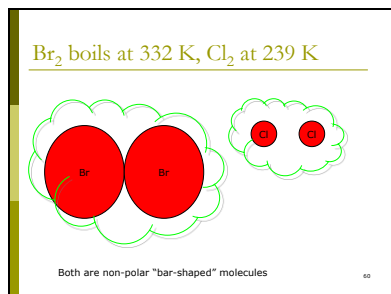
This is a warning...your final warning...

There are limits to the easy comparisons.

If you have two structurally similar molecules, then the heavier one will have the higher boiling point.

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Polar molecules are higher...

Permanent dipoles are sort of a bonus.

Take two similarly shaped molecules with similar molar masses and the polar one will have a higher boiling point than the lower one.

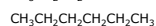
But if the molar masses are different enough, the polar nature won't save you.

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A few molecules

Molecule	Molar mass	Polar/ non-polar	Boiling point
chloropropane	78.5 g/mol	Weakly Polar	320 K
Hexane	86 g/mol	Non-polar	342 K
Chlorine	70.9 g/mol	Non-polar	239 K
Calcium sulfide	72.1 g/mol	ionic	Melts at 2800 K
Sulfur dioxide	64 g/mol	Polar	263 K
Water	18.02 g/mol	Polar	373 K



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So if you are comparing 2 molecules:

1. Look for ionic compounds – they have the strongest forces – trumps EVERYTHING
2. Look for hydrogen bonding – hydrogen bonding is the 2nd strongest and will usually swamp van der Waal's if the molecules are SIMILAR size
3. Van der Waal's forces – heavier wins
4. Dipole-dipole forces – sort of a tie-breaker

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Limits of hydrogen bonding:

Octane – $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$
Molar mass = 114.23 g/mol
Non-polar molecule
Boiling point = 399 K

Ethanol – $\text{CH}_3\text{CH}_2\text{-OH}$
Molar mass = 46.07 g/mol
Hydrogen Bonding
Boiling point = 351 K

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Limits of hydrogen bonding:

Octane – $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$
Molar mass = 114.23 g/mol
Non-polar molecule
Boiling point = 399 K

Ethanol – $\text{CH}_3\text{CH}_2\text{-OH}$
Molar mass = 46.07 g/mol
Hydrogen Bonding
Boiling point = 351 K

Water – H_2O
Molar mass = 18.02 g/mol
Hydrogen bonding
Boiling point = 373 K

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Limits of hydrogen bonding:

Octane – $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$
Molar mass = 114.23 g/mol
Non-polar molecule
Boiling point = 399 K

Octanol – $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$
Molar mass = 130.23 g/mol
Hydrogen bonding
Boiling point = 468 K

Bigger

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Limits of hydrogen bonding:

Ethanol – $\text{CH}_3\text{CH}_2\text{OH}$
Molar mass = 46.07 g/mol
Hydrogen Bonding
Boiling point = 351 K

Ethane – CH_3CH_3
Molar mass = 30.07 g/mol
Non-polar
Boiling point = 185 K

Almost double! The hydrogen bond is a much bigger part of the smaller molecule

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Bottom Line

The more similar the molecules are in size and shape the easier it is to determine the size of the relative forces.

If they are very different in size (ethanol vs. octane) or shape (HBr vs. Cl_2) we are just making educated guesses.

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In order of importance

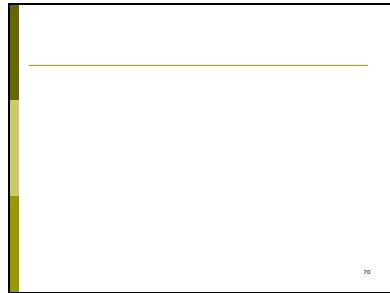
1. Ionic forces (biggest by a lot)
2. Hydrogen bonding (special case of...)
3. Dipole-Dipole
4. Van der Waal's

But 3 and 4 are much weaker than 1 and 2.

3 only matters if the molecules are similar sizes.

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Here's some...

Physical properties that show "intermolecular forces":

1. Boiling point
2. Melting point
3. Surface tension
4. Viscosity
5. Capillary action
6. Evaporation

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Phase changes

Intermolecular Forces are attractions between molecules.

Temperature is a measure of kinetic energy.

Boiling Point (or Freezing Point) are measures of the strength of intermolecular forces: the higher the temperature, the more kinetic energy required to separate the molecules.

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Not just temperature...

We mentioned TWO things that affected molecules and their interactions:

1. Energy
2. Space

Another way of looking at "space" is pressure.

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What is "pressure"?

Pressure = $\frac{\text{Force}}{\text{Area}}$

Pressure is squeezing the molecules together!

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Phase Changes

You can create a phase change, by changing the temperature.

Consider a flask full of steam at 200°C.

If I start cooling it down, what happens?

It condenses into liquid water. When?

NOT (necessarily) 100°C.

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Normal Boiling Point

100°C is the "normal boiling point" of water.
What's the "normal" for?

Normal means at standard pressure, 1 atm.

One way to condense steam is to decrease the temperature, another way is to increase the pressure.

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It's all about forces!

Pressure

Intermolecular Force

Kinetic Energy

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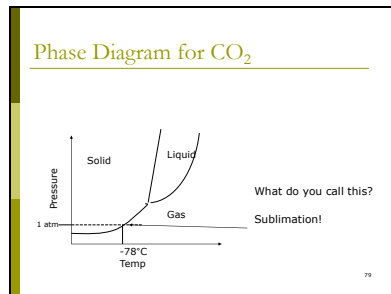
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Phase Diagrams

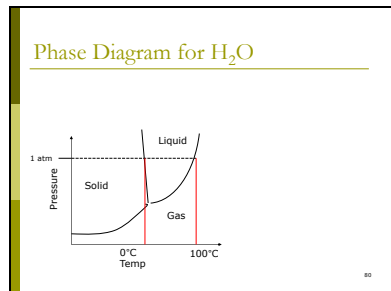
A "phase diagram" collects all the P, T and phase information and displays it in one simple graph.

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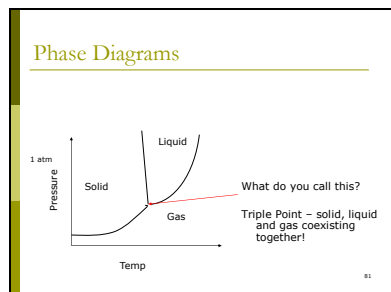
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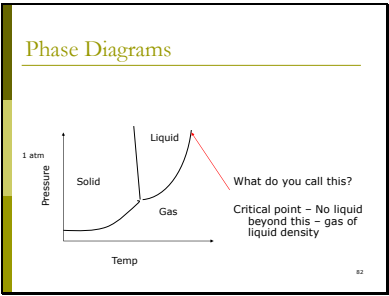
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Energy of Phase Changes

How do you define "boiling"?

Vapor pressure = atmospheric pressure

What's vapor pressure?

It's the pressure exerted by the vapor above a liquid.

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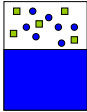
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As you raise T , you raise P_{vap} until $P_{\text{vap}} = P_{\text{atm}}$

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As you raise T , you raise P_{vap} until $P_{\text{vap}} = P_{\text{atm}}$

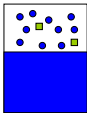


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Remember $P_{\text{tot}} = P_1 + P_2 +$

The vapor crowds out the air above the solution since P_{tot} must always be P_{atm}

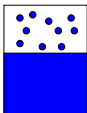


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Remember $P_{\text{tot}} = P_1 + P_2 +$

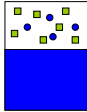
P_{tot} must always be P_{atm} . When $P_{\text{vap}} = P_{\text{atm}}$, it's all water vapor and WE ARE BOILING!



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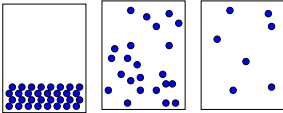
What do you have to do to become "vapor"?
You have to go from a liquid to a gas!
What do you need to do to go from a liquid to a gas?
GAIN ENERGY!



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Remember, all molecules like each other.
So the difference between a solid, a liquid
and a gas...



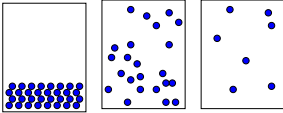
Solid Liquid Gas

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...is all relative to the Energy

There are two different energies (or forces).
The attraction between molecules, the
individual energy of the molecules.



Solid Liquid Gas

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Suppose I tie myself to one of you using a noodle. Could you escape?

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Of course you could.

You just start walking away and the noodle breaks.

Suppose I tie myself to you using a piece of thread?

You may have to walk faster or pull harder but you can still break away.

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Suppose I tie myself to you using a piece of copper wire?

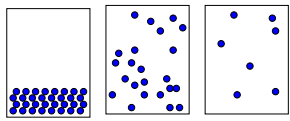
You may have to run or tug or get your friends to also tug, but you can break the wire.

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Same for phases of matter.

They like each other, you want to separate them
you need to overcome the "like". Easiest way:
heat 'em up so they are moving faster!



Solid Liquid Gas

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Making a phase change...

Suppose I start with 100 g of ice at -40°C (1 atm)
and start heating it up, what happens?

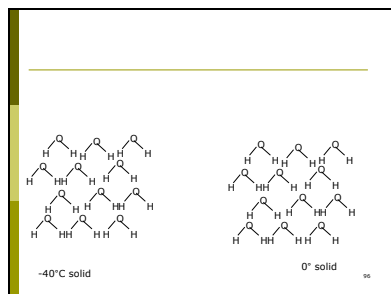
The ice gets warmer and warmer until...melting
point!

Suppose I am ice at 0°C , do I just spontaneously
melt?

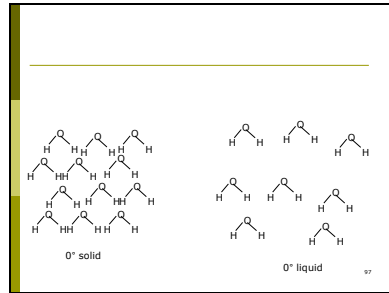
Not exactly. I am warm enough, but I'm still a
solid and my molecules are still "associated" with
each other. I need to get ripped away from my
brothers.

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At the phase transition temperature...

...you still need energy to make the transition.

Going from solid to liquid, this is called the "heat of fusion" (ΔH_{fus})

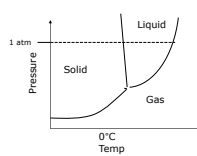
Going from liquid to gas, this is called "heat of vaporization" ($\Delta H_{\text{vap}}^\circ$)

Going from solid to gas, this is called the "heat of sublimation" ($\Delta H_{\text{sub}}^\circ$)

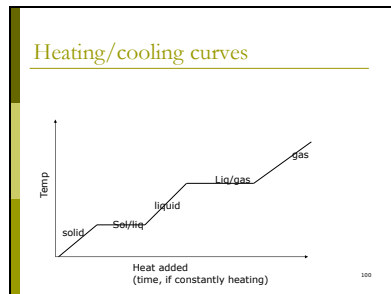
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Phase Diagram for H₂O

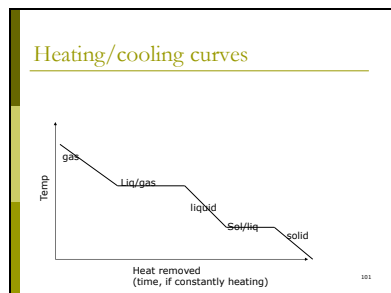
The "°" in the ΔH° means standard conditions.



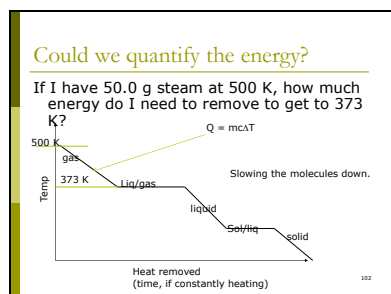
Slide 100



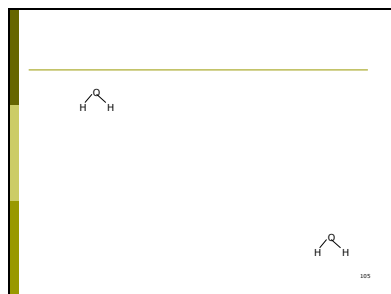
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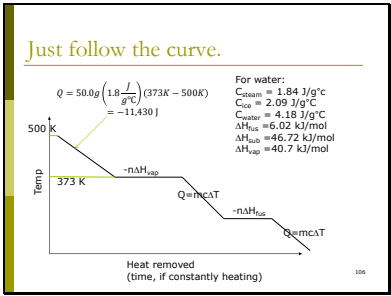
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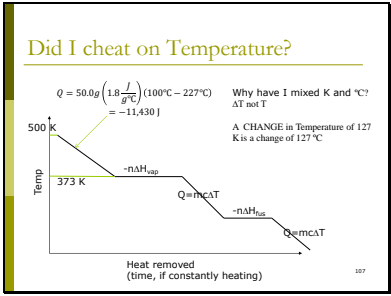
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[illegible]

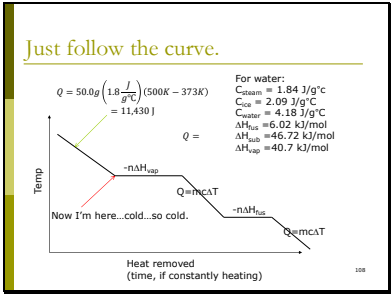
Slide 106



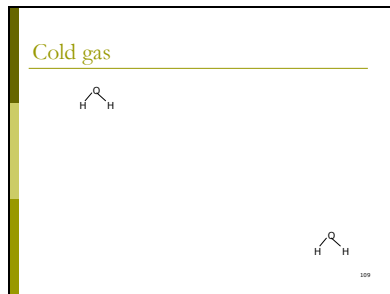
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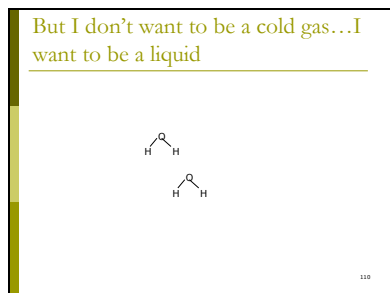
Slide 108



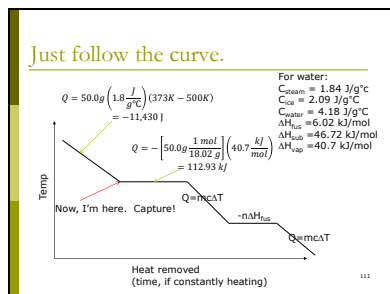
Slide 109



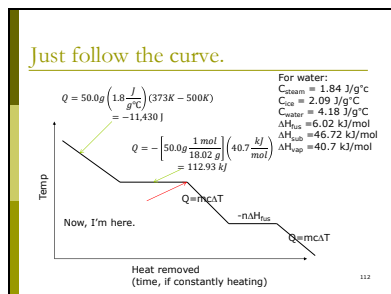
Slide 110



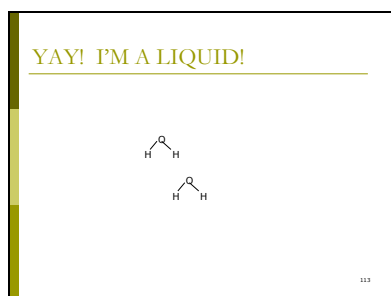
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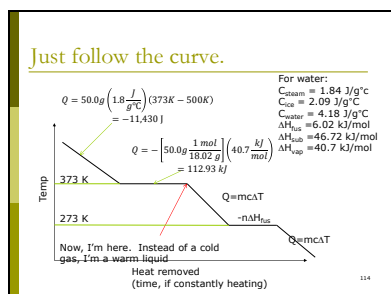
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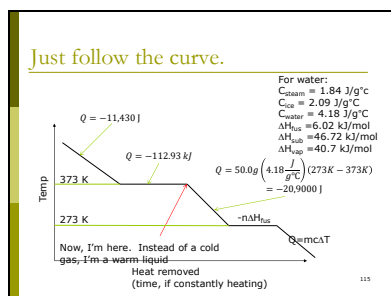
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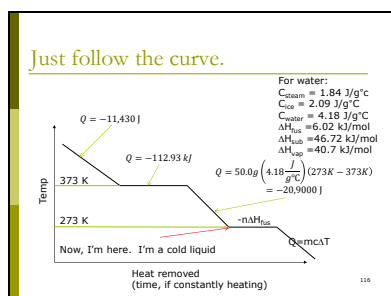
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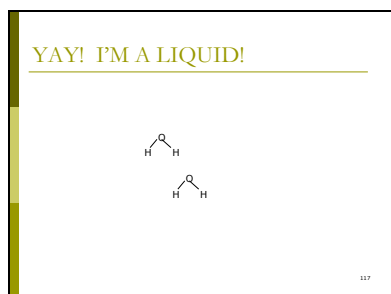
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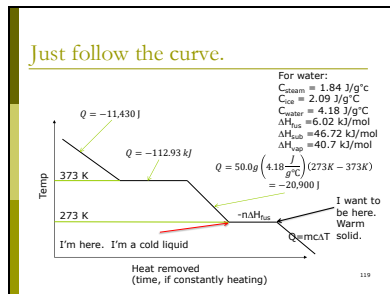
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I don't want to be a cold liquid. I want to be a warm solid!

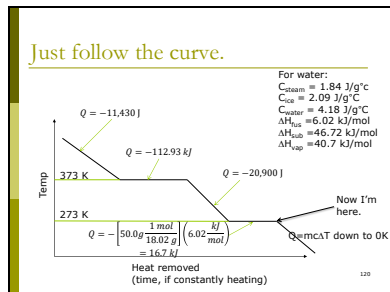
0° solid

0° liquid

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A little problem

I have 50 g of ice at 100 K. How much energy would I need to add to get steam at 500 K?

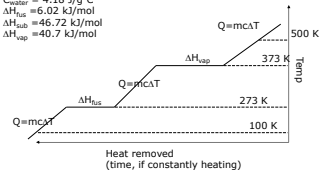
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Just follow the curve.

For water:

C_{steam}	$= 1.84 \text{ J/g}^\circ\text{C}$
C_{ice}	$= 2.09 \text{ J/g}^\circ\text{C}$
C_{water}	$= 4.18 \text{ J/g}^\circ\text{C}$
ΔH_{fus}	$= 6.02 \text{ kJ/mol}$
ΔH_{sub}	$= 46.72 \text{ kJ/mol}$
ΔH_{vap}	$= 40.7 \text{ kJ/mol}$



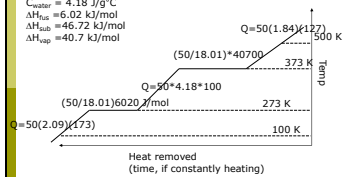
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Just follow the curve.

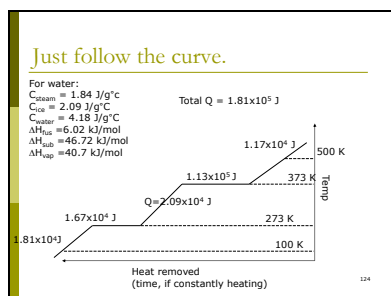
For water:

$C_{\text{steam}} = 1.84 \text{ J/g}^\circ\text{C}$
 $C_{\text{ice}} = 2.09 \text{ J/g}^\circ\text{C}$
 $C_{\text{water}} = 4.18 \text{ J/g}^\circ\text{C}$
 $\Delta H_{\text{fus}} = 6.02 \text{ kJ/mol}$
 $\Delta H_{\text{sub}} = 46.72 \text{ kJ/mol}$
 $\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$

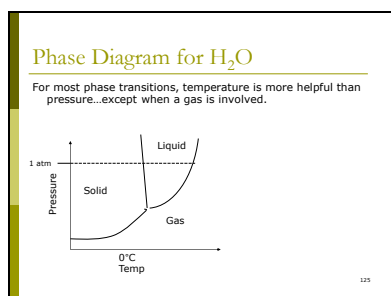


123

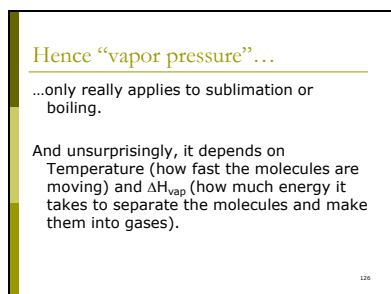
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Vapor Pressure

Vapor pressure depends on temperature. Vapor pressure also depends on ΔH_{vap}

Clausius-Clapeyron equation:

$$\ln P_{\text{vap}} = \frac{-\Delta H_{\text{vap}}}{R} \times \frac{1}{T} + C$$

Where C is a constant, R is the ideal gas constant.

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Not completely useful in this form

Clausius-Clapeyron equation:

$$\ln P_{\text{vap}} = \frac{-\Delta H_{\text{vap}}}{R} \times \frac{1}{T} + C$$

If I want to calculate P_{vap} , I need to know ΔH_{vap} , C, and T. Except for T, the other two parameters are specific to each compound measured. But math (as ALWAYS!) can Save The Day!!

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Why did I write it that way?

Clausius-Clapeyron equation:

$$\ln P_{\text{vap}} = \frac{-\Delta H_{\text{vap}}}{R} \times \frac{1}{T} + C$$

I could have just written it as:

$$\ln P_{\text{vap}} = \frac{-\Delta H_{\text{vap}}}{RT} + C$$

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Why did I write it that way?

Clausius-Clapeyron equation:

$$\ln P_{\text{vap}} = \frac{-\Delta H_{\text{vap}}}{R} \times \frac{1}{T} + C$$

Looks like:

$$y = mx + b$$

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If I'm doing an experiment

Clausius-Clapeyron equation:

$$\ln P_{\text{vap}} = \frac{-\Delta H_{\text{vap}}}{R} \times \frac{1}{T} + C$$

If I plot $\ln P_{\text{vap}}$ vs. $\frac{1}{T}$ I should get a straight line with a y-intercept of C and a slope of $\frac{-\Delta H_{\text{vap}}}{R}$

This is how you would find C and ΔH_{vap}

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There's also the short cut

Maybe you don't want to do the whole experiment! And maybe someone else has already determined ΔH_{vap} (you did the enthalpy lab!)

Algebra is your BESTEST friend!

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Vapor Pressure

And then there's "normal":

$$\ln \frac{P_{\text{vap},1}}{P_{\text{vap},2}} = \frac{-\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

I usually know the "normal boiling point" of a material...which is?

The boiling point at $P_{\text{atm}} = 1 \text{ atm}$. Since boiling occurs when $P_{\text{vap}} = P_{\text{atm}}$, I know one set of P_{vap} and T !

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Sample problem:

What is the vapor pressure of water at 50°C?

I say vapor pressure, you think...
Clausius-Clapeyron!

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Vapor Pressure

$$\ln \frac{P_{\text{vap},1}}{P_{\text{vap},2}} = \frac{-\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

What do I know?

$P_{\text{vap},1} = ?$
 $P_{\text{vap},2} = ?$
 $\Delta H_{\text{vap, water}} = ?$
 $T_2 = ?$
 $T_1 = ?$
 $R = ?$

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Vapor Pressure

$$\ln \frac{P_{vap,1}}{P_{vap,2}} = \frac{-\Delta H_{vap}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

What do I know?

$$P_{\text{vap}1} = 1 \text{ atm}$$
$$P_{\text{vap}2} = ?$$
$$\Delta H_{\text{vap, water}} = 40.7 \text{ kJ/mol at boiling point (pg 472, Tro)}$$
$$= 44.0 \text{ kJ/mol at } 25^\circ\text{C}$$
$$T = 50^{\circ}\text{C} = 323.15\text{ K}$$
$$T_2 = 50^\circ\text{C} = 323.15\text{ K}$$
$$T_1 = 100\text{ }^{\circ}\text{C} = 373.15\text{ K}$$
$$R = 8.314 \text{ J/(mol K)}$$

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Plugging and chugging time...

$$\ln \frac{P_{vap,1}}{P_{vap,2}} = \frac{-\Delta H_{vap}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

What do I know?

$$P_{\text{vap}1} = 1 \text{ atm}$$
$$P_{\text{vap}2} = ?$$
$$\Delta H_{\text{vap, water}} = 44.0 \text{ kJ/mol at } 25^\circ\text{C}$$
$$T_2 = 50^\circ\text{C} = 323.15\text{ K}$$
$$T_2 = 50^\circ\text{C} = 323.15\text{ K}$$
$$T_1 = 100^\circ\text{C} = 373.15\text{ K}$$
$$T_1 = 100^\circ\text{C} = 373.15\text{ K}$$
$$R = 8.314\text{ J/(mol K)}$$

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Plugging and chugging time...

$$\ln \frac{1 \text{ atm}}{P_{\text{vap},2}} = \frac{-44.0 \times 10^3 \frac{\text{J}}{\text{mol}}}{8.314 \frac{\text{J}}{\text{mol} \times \text{K}}} \left(\frac{1}{373.15} - \frac{1}{323.15} \right)$$

Whatever you do, DON'T ROUND!

$$\ln \frac{1 \text{ atm}}{P_{\text{vap},2}} = \frac{-44.0 \times 10^3 \frac{\text{J}}{\text{mol}}}{8.314 \frac{\text{J}}{\text{mol} \times \text{K}}} (0.0026798 - 0.00309453)$$

$$\ln\left(\frac{1 \text{ atm}}{p_{\text{vap},2}}\right) = 2.194446789$$

How do I isolate P_{vap2} ?

That's right e^x !

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Plugging and chugging time...

$$\ln\left(\frac{1 \text{ atm}}{P_{\text{vap},2}}\right) = 2.194446789$$

$$\frac{1 \text{ atm}}{P_{\text{vap},2}} = e^{2.194446789} = 8.97880$$

$$P_{vap,2} = \frac{1 \text{ atm}}{8.97880} = 0.111373 \text{ atm}$$

Does this make sense?
It is less than 1 atm and I'm below the boiling point!

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Another little problem

What is the boiling point of water at the top of Mt. Everest where the average atmospheric pressure is 0.64 atm?

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Vapor Pressure

$$\ln \frac{P_{vap,1}}{P_{vap,2}} = \frac{-\Delta H_{vap}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

What do I know?

$$P_{\text{vap}1} = ?$$
$$P_{\text{vap}2} = ?$$
$$\Delta H_{\text{vap, water}} = ?$$
$$T_2 = ?$$
$$T_1 = ?$$
$$R = ?$$

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Vapor Pressure

$$\ln \frac{P_{\text{vap},1}}{P_{\text{vap},2}} = \frac{-\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

What do I know?

$P_{\text{vap},1} = 1 \text{ atm}$
 $P_{\text{vap},2} = 0.64 \text{ atm}$
 $\Delta H_{\text{vap, water}} = 44.0 \text{ kJ/mol at } 25^\circ\text{C}$
 $T_2 = ?$
 $T_1 = 100^\circ\text{C} = 373.15 \text{ K}$
 $R = 8.314 \text{ J/(mol K)}$

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Plugging and chugging time...

$$\ln \frac{1 \text{ atm}}{0.64 \text{ atm}} = \frac{-44.0 \times 10^3 \frac{\text{J}}{\text{mol}}}{8.314 \frac{\text{J}}{\text{mol} \times \text{K}}} \left(\frac{1}{373.15 \text{ K}} - \frac{1}{T_2} \right)$$

Whatever you do, DON'T ROUND!

$\ln 1.5625 = -5292.278 \quad [0.0026798 - 1/T_2]$

$0.446287 = -5292.278 [0.0026798 - 1/T_2]$
 $-0.00008432797 = 0.0026798 - 1/T_2$
 $1/T_2 = 0.0027642$
 $T_2 = 361.77 \text{ K} = 88.6^\circ\text{C}$
 Does this make sense?
 Lower atmospheric pressure, lower boiling point!

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