

Slide 1

Intermolecular Forces

Love & Hate in the Molecular Realm

1

Slide 2

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

2

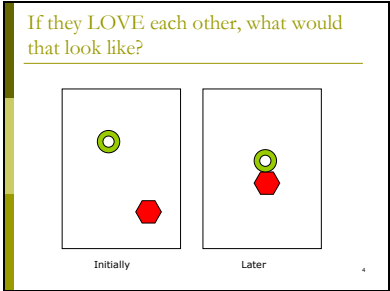
Slide 3

**REMEMBER:
MOLECULES MOVE!**

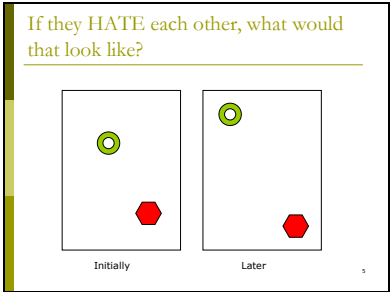
(except at 0 K)

3

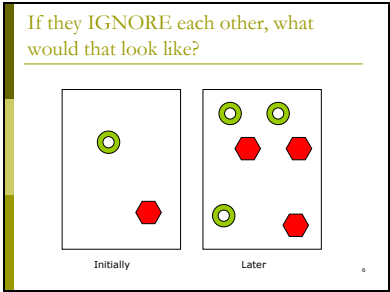
Slide 4



Slide 5



Slide 6

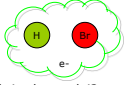


Slide 7

What determines LOVE or HATE?

The structure of the molecule.

What is the structure of a molecule?



What's in the nuclei?
Protons!

7

Slide 8

Molecular structure is all about...

POSITIVE & NEGATIVE CHARGES!

So Love & Hate is all about...

Opposites attract, like repel!

8

Slide 9

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

9

Slide 10

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

10

Slide 11

Instantaneous dipole-induced dipole forces

The electron cloud is mobile. Charge density is constantly moving around

11

Slide 12

How Great is THAT!?!?

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

12

Slide 13

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!

13

Slide 14

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.

14

Slide 15

Alkanes

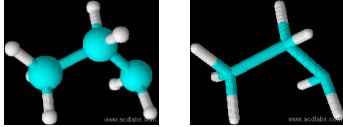
Methane - CH₄
Ethane - CH₃CH₃
Propane - CH₃CH₂CH₃
Butane - CH₃CH₂CH₂CH₃
Pentane - CH₃CH₂CH₂CH₂CH₃
Hexane - CH₃CH₂CH₂CH₂CH₂CH₃
Heptane - CH₃CH₂CH₂CH₂CH₂CH₂CH₃
Octane - CH₃CH₂CH₂CH₂CH₂CH₂CH₂CH₃

15

Slide 16

What do they look like?

Here's propane (C_3H_8): $CH_3CH_2CH_3$

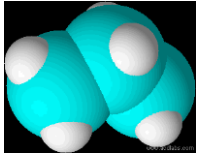


16

Slide 17

When you include the electrons...

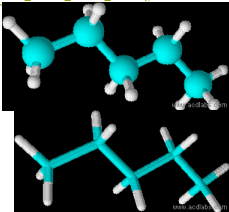
"Space-filling model" $CH_3CH_2CH_3$



17

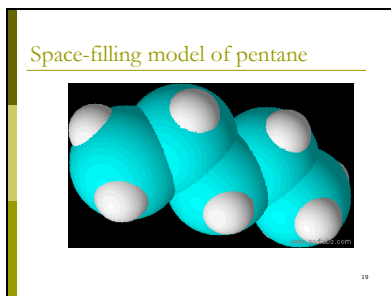
Slide 18

Here's pentane (C_5H_{12}):
 $CH_3CH_2CH_2CH_2CH_3$



18

Slide 19



Slide 20

Pentane is just a "longer caterpillar" than propane.

That makes it easier to compare these molecules, they are homologues.

20

Slide 21

What do you know about these molecules?

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 - $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$

21

Slide 22

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

22

Slide 23

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

23

Slide 24

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.

24

Slide 25

ALL MOLECULES
MOVE!

(except at 0 K)

25

Slide 26

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...

26

Slide 27

Solids, Liquids, and Gases & Heat

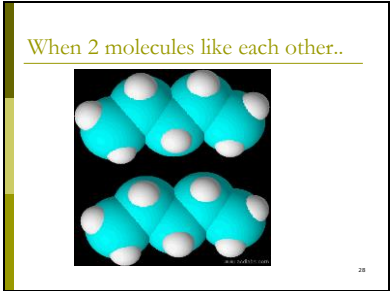
What happens when you heat up a solid?

Eventually it melts - why?

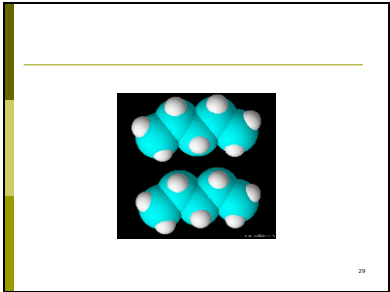
Adding heat adds energy to the molecules, when they have enough energy they can escape their attraction to their neighbors!

27

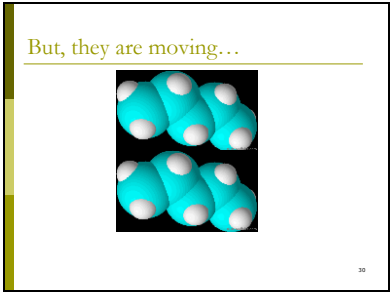
Slide 28



Slide 29

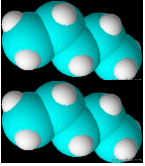


Slide 30



Slide 31

If they are HOT HOT HOT!



31

Slide 32

If they are very COLLLDDDD



32

Slide 33

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.]

33

Slide 34

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.

34

Slide 35

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.

35

Slide 36

Dipole - Dipole Interactions

Permanent dipole

Permanent dipole

36

Slide 37

Seems **REPULSIVE** – but it's really not

Permanent dipole

Permanent dipole

MOLECULES ARE MOBILE. They always align themselves. That's why I say that in the end all intermolecular forces are attractive.

37

Slide 38

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.

38

Slide 39

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**.

39

Slide 40

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.

40

Slide 41

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

Strongly polar dipole

Strongly polar dipole

41

Slide 42

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.

42

Slide 43

Compare H₂O to H₂S

The boiling point of water is 373 K.

The boiling point of H₂S is 213 K.

H₂S is a gas at room temperature while water is a liquid!

No FON, no Hydrogen bonding

43

Slide 44

Ion – Ion Interactions

Actual separation of charge

Actual separation of charge

44

Slide 45

Ion-Ion interactions

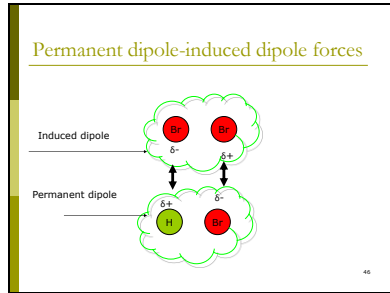
The strongest possible interaction.

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

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

45

Slide 46



Slide 47

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.

47

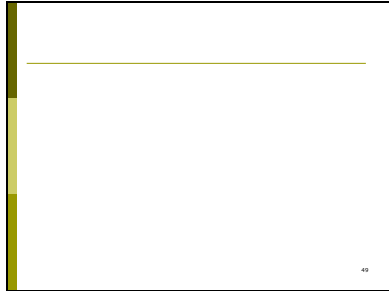
Slide 48

All the forces...

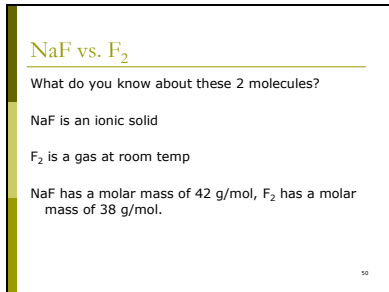
1. Van der Waal's/Dispersion forces – FIRST consideration. Weakest for single bond BUT it is a more global force. Heavier molecules have bigger VDW forces.
2. Dipole-Dipole forces – add on to VDW forces (with ONE exception - #3). If the molecules have similar mass and shape. The one with a permanent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

48

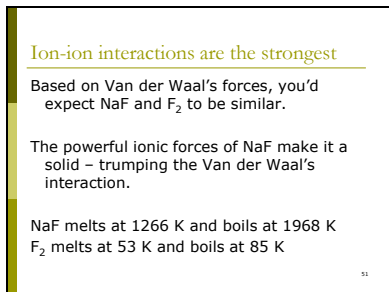
Slide 49



Slide 50



Slide 51



Slide 52

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

52

Slide 53

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...

53

Slide 54

HBr boils at 207 K, Cl₂ boils at 239 K

WTWikipedia?!?!?

54

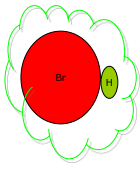
Slide 55

So, why doesn't it?
Geometry is also an issue!

55

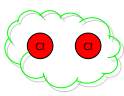
Slide 56

Geometry



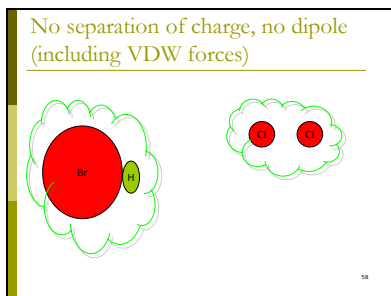
56

Slide 57



57

Slide 58



Slide 59

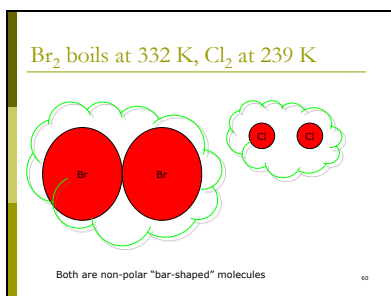
This is a warning...your final warning...

There are limits to the easy comparisons.

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

59

Slide 60



Slide 61

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.

61

Slide 62

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

CH3CH2CH2Cl
CH3CH2CH2CH2CH2CH3

62

Slide 63

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

63

Slide 64

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

64

Slide 65

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

65

Slide 66

Limits of hydrogen bonding:

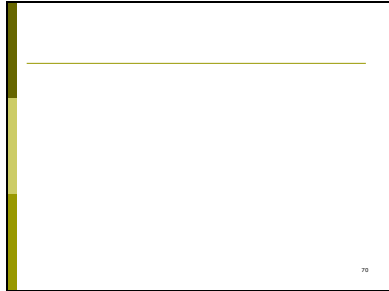
Octane – $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$
Molar mass = 114.23 g/mol
Non-polar molecule
Boiling point = 399 K

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

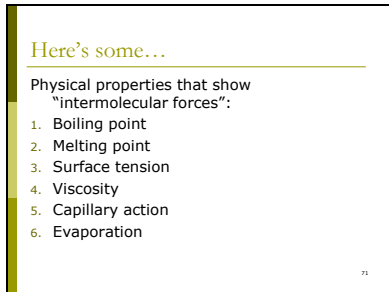
Bigger

66

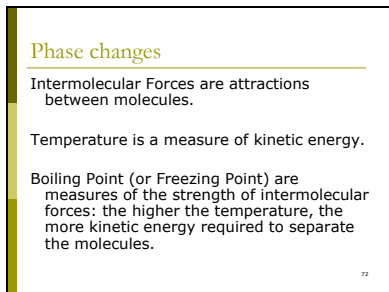
Slide 70



Slide 71



Slide 72



Slide 73

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.

73

Slide 74

What is "pressure"?

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

Pressure is squeezing the molecules together!

74

Slide 75

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.

75

Slide 76

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.

76

Slide 77

It's all about forces!

Pressure

Kinetic Energy

Intermolecular Force

77

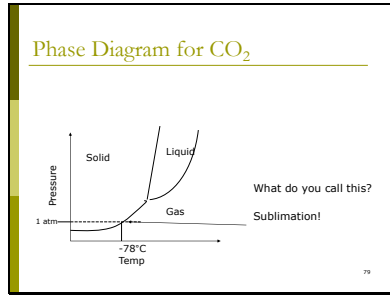
Slide 78

Phase Diagrams

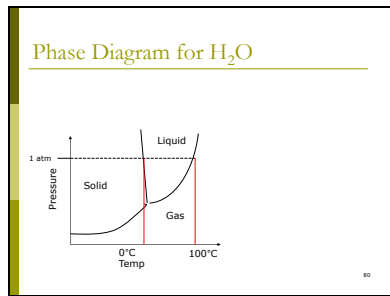
A "phase diagram" collects all the P, T and phase information and displays it in one simple graph.

78

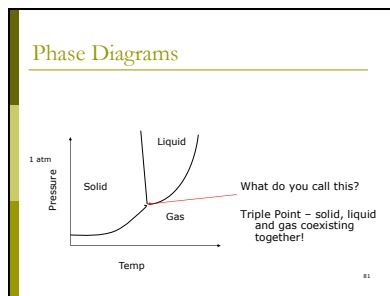
Slide 79



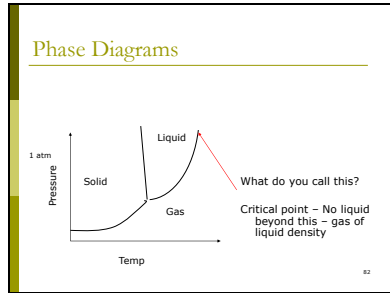
Slide 80



Slide 81



Slide 82



Slide 83

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.

83

Slide 84

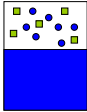
As you raise T, you raise P_{vap} until $P_{\text{vap}} = P_{\text{atm}}$

A diagram showing a blue rectangular container. Above the liquid surface, there are several small squares and circles representing gas molecules.

84

Slide 85

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

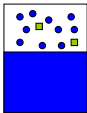


85

Slide 86

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

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

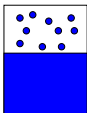


86

Slide 87

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

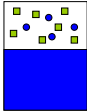
P_{tot} must always be P_{atm} . When $P_{\text{vap}} = P_{\text{atm}}$, it's all water vapor and WE ARE BOILING!



87

Slide 88

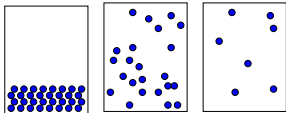
What do you have to do to become "vapor"?
You have to go from a liquid to a gas!
What do you need to do to go from a liquid to a gas?
GAIN ENERGY!



88

Slide 89

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



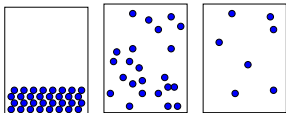
Solid Liquid Gas

89

Slide 90

...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

90

Slide 91

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

91

Slide 92

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.

92

Slide 93

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

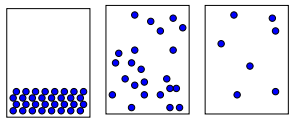
You may have to run or tug or get your friends to also tug, but you can break the wire.

93

Slide 94

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

94

Slide 95

Making a phase change...

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

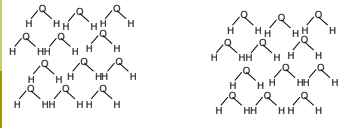
The ice gets warmer and warmer until...melting point!

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.

95

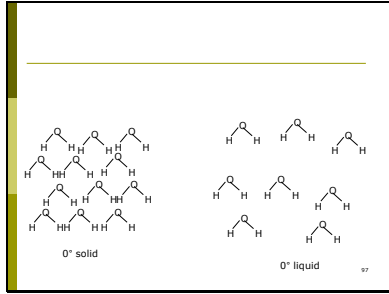
Slide 96



-40°C solid 0° solid

96

Slide 97



Slide 98

At the phase transition temperature...
...you still need energy to make the transition.

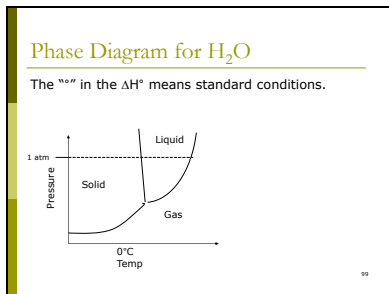
Going from solid to liquid, this is called "heat of fusion" ($\Delta H_{\text{fus}}^\circ$)

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

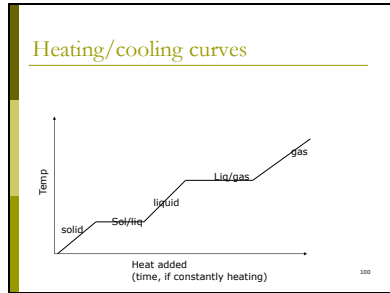
Going from solid to gas, this is called the "heat of sublimation" ($\Delta H_{\text{sub}}^\circ$)

98

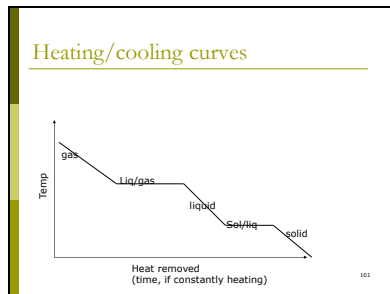
Slide 99



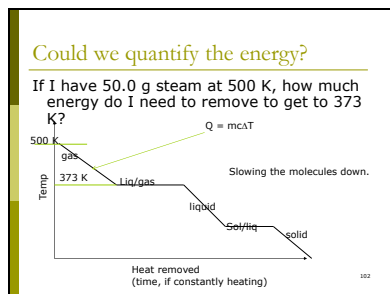
Slide 100



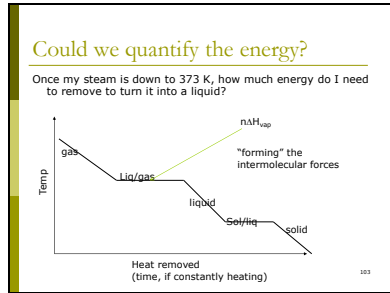
Slide 101



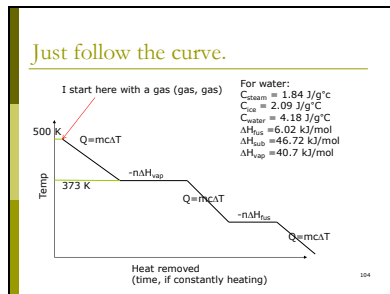
Slide 102



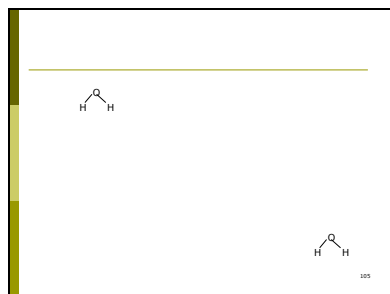
Slide 103



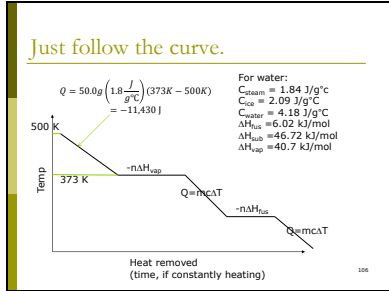
Slide 104



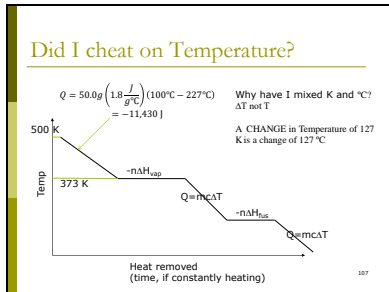
Slide 105



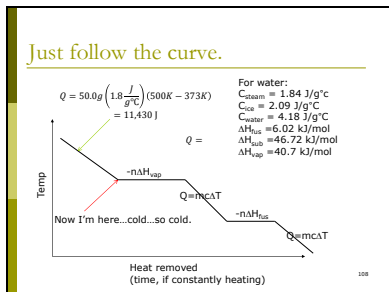
Slide 106



Slide 107




Slide 108



Slide 109


Cold gas



109

Slide 110

But I don't want to be a cold gas... I want to be a liquid



110

Slide 111

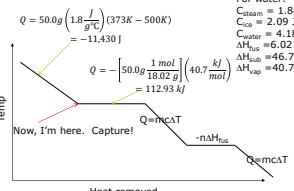
Just follow the curve.

For water:

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

$Q = 50.0\text{g} \left(1.8 \frac{\text{J}}{\text{g}^\circ\text{C}} \right) (373\text{K} - 500\text{K})$
 $= -11,430 \text{ J}$

$Q = - \left[50.0\text{g} \frac{1 \text{ mol}}{18.02 \text{ g}} \right] \left(40.7 \frac{\text{kJ}}{\text{mol}} \right)$
 $= 112.93 \text{ kJ}$



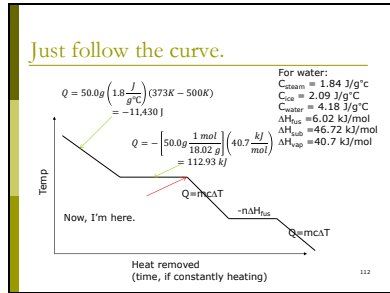
Temp

Now, I'm here. Capture!

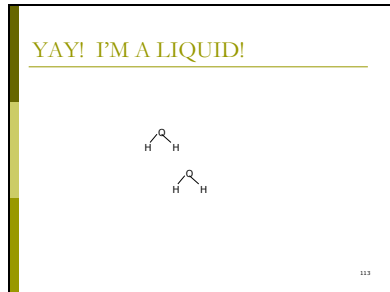
Heat removed (time, if constantly heating)

111

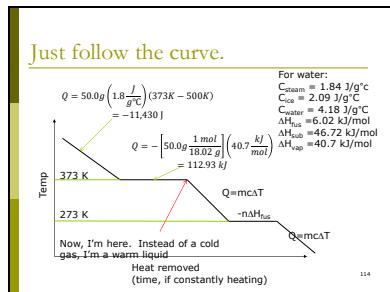
Slide 112



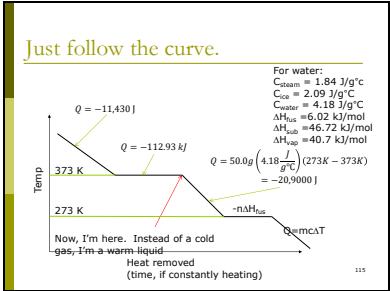
Slide 113



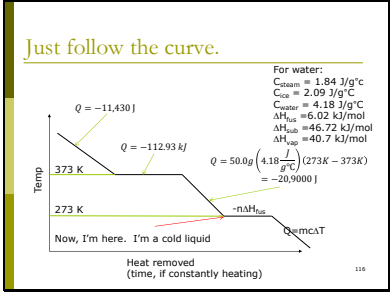
Slide 114



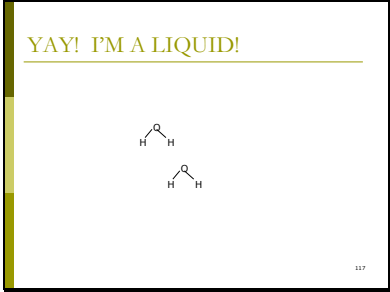
Slide 115



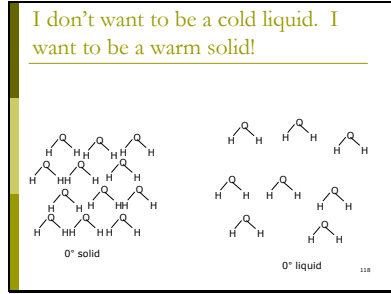
Slide 116



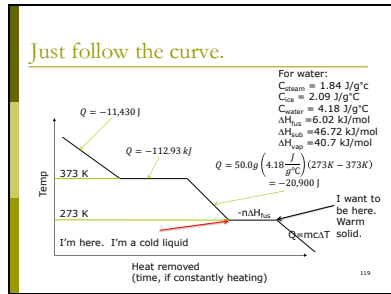
Slide 117



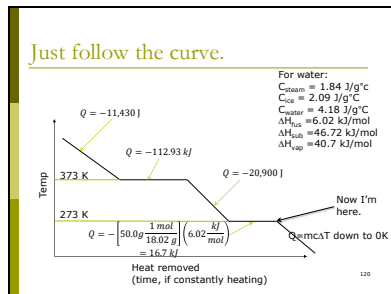
Slide 118



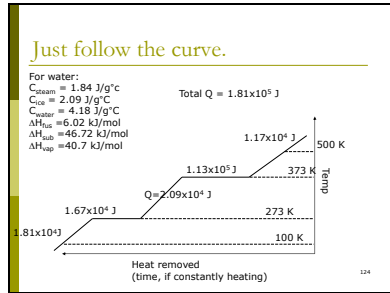
Slide 119



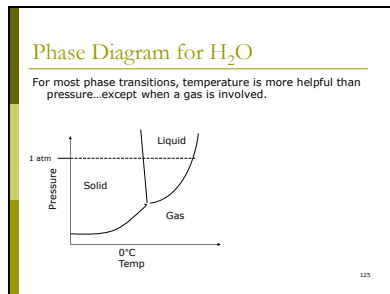
Slide 120



Slide 124



Slide 125



Slide 126

Hence "vapor pressure"...

...only really applies to sublimation or boiling.

And unsurprisingly, it depends on Temperature (how fast the molecules are moving) and ΔH_{vap} (how much energy it takes to separate the molecules and make them into gases).

126

Slide 127

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.

127

Slide 128

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!!

128

Slide 129

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:

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

129

Slide 130

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:

$$y = mx + b$$

130

Slide 131

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}

131

Slide 132

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!

132

Slide 136

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!

136

Slide 137

Sample problem:

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

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

137

Slide 138

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?

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

138

Slide 142

Plugging and chugging time...

$$\ln\left(\frac{1 \text{ atm}}{P_{\text{vap},2}}\right) = 2.194446789$$
$$\frac{1 \text{ atm}}{P_{\text{vap},2}} = e^{2.194446789} = 8.97880$$
$$P_{\text{vap},2} = \frac{1 \text{ atm}}{8.97880} = 0.111373 \text{ atm}$$

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

142

Slide 143

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?

143

Slide 144

Vapor Pressure

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

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

144

Slide 145

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

145

Slide 146

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!
 $\ln 1.5625 = -5292.278 \quad [0.0026798 - 1/T_2]$
 $0.446287 = -5292.278 \quad [0.0026798 - 1/T_2]$
 $-0.00008432797 = 0.0026798 - 1/T_2$
 $1/T_2 = 0.0027642$
 $T_2 = 361.77 \text{ K} = 88.6^\circ\text{C}$
Does this make sense?
Lower atmospheric pressure, lower boiling point! 146
