

Chemistry 104 Chapter Seven PowerPoint Notes

Gases, Liquids and Solids Chapter 7

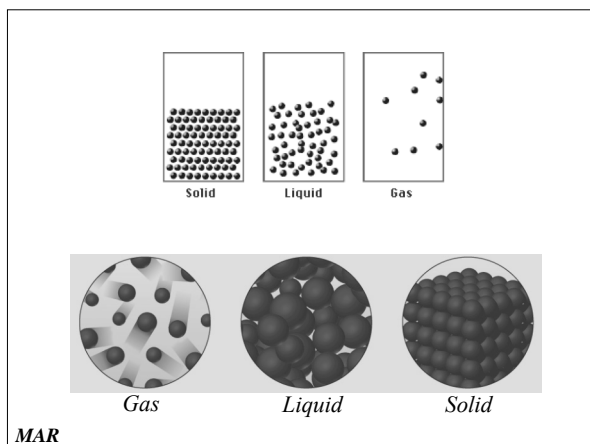
Chemistry 104
Professor Michael Russell

Phases of Matter

Matter can exist in any of three phases - solid, liquid, and gas - depending on the **strength** of the **attractive forces** among the particles in the substance.

- **Solids:** Strongest attractive forces; atoms, molecules, or ions held together in a specific arrangement.
- **Liquids:** Strong attractive forces which pull particles close together but still allow considerable freedom to move.
- **Gases:** Weak attractive forces, particles move about freely and are far from each other.

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Transformation of a substance from one state to another referred to as **phase changes** or **changes of state**.

Changes of state *reversible* and characterized by an **energy change**.

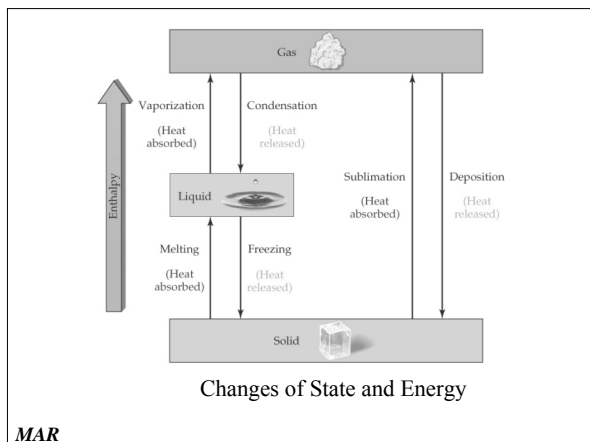
Melting point (mp): Temperature at which solid and liquid are in *equilibrium*.

Boiling point (bp): Temperature at which liquid and gas are in equilibrium.

Sublimation: Process whereby a solid changes directly to a gas without being a liquid.

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Gases and the Kinetic Molecular Theory

The behavior of gases can be explained by the **kinetic molecular theory (KMT)**. Principles of the KMT are:

1. Gases consist of many particles moving at random with no attractive forces between them. Random motion allows different gases to mix quickly.
2. Most of gas volume is *empty space* which accounts for easy compression and low densities.
3. The average energy of gas particles proportional to the Kelvin temperature.
4. The pressure of a gas is the result of collisions of gas particles with the walls of container.
5. Gases that obey all the assumptions of KMT called

MAR **ideal gases**.

P, V and T

Pressure (P) important for gases.

The *atmosphere* is a blanket of air pressing down on us at all times.

Atmospheric pressure not constant.

Units of pressure: *millimeters of mercury* (mm Hg) and often called a *torr*. SI unit of pressure is *pascal* (Pa).

$$1 \text{ mm Hg} = 1 \text{ torr} = 133.32 \text{ Pa}$$

$$1 \text{ atm} = 760 \text{ mm Hg} = 101,325 \text{ Pa}$$

Volume (V) important, usually in L or mL

Temperature (T) important, *has to be Kelvin (K)*

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Boyle's Law: The Relation Between Volume and Pressure

Gas laws make it possible to predict the influence of pressure (P), volume (V), temperature (T) and moles (n) of any gas or mixture of gases.

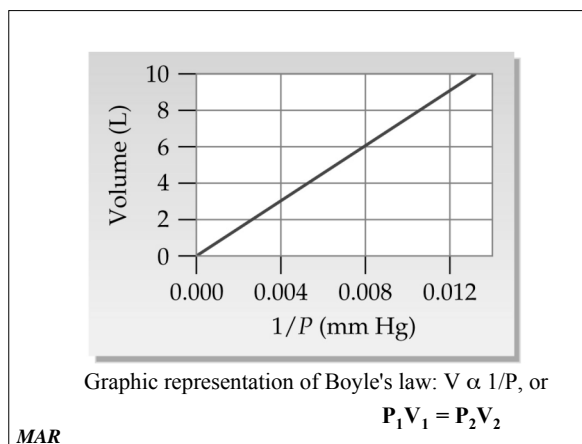
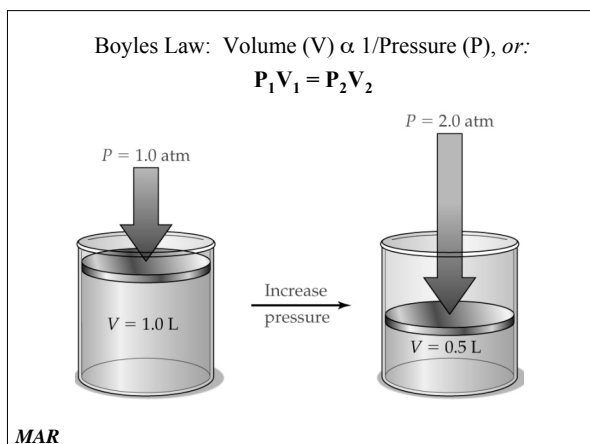
Boyles' law: "*The volume of a fixed amount of gas at a constant temperature is inversely proportional to its pressure.*"

That is, the volume and pressure change in opposite directions.

As pressure goes up, volume goes down; as pressure goes down, volume goes up.

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Boyle's Law: Example

What will be the pressure of a gas that has its volume changed from 5.0 L to 10.0 L if the original pressure is 2.5 atm?

Use: $P_1V_1 = P_2V_2$

Given:

$$\begin{aligned} V_1 &= 5.0 \text{ L} \\ V_2 &= 10.0 \text{ L} \\ P_1 &= 2.5 \text{ atm} \end{aligned}$$

Boyle's Law: Example

Solve for P_2

$$P_2 = \frac{P_1V_1}{V_2}$$

Substitute in the values

$$\begin{aligned} P_2 &= \frac{(2.5 \text{ atm}) (5.0 \cancel{\text{L}})}{10.0 \cancel{\text{L}}} \\ P_2 &= 1.3 \text{ atm} \end{aligned}$$

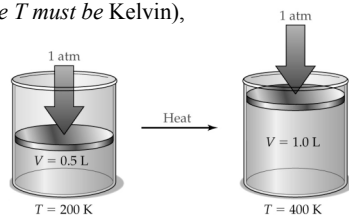
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Charles's Law: The Relation Between Volume and Temperature

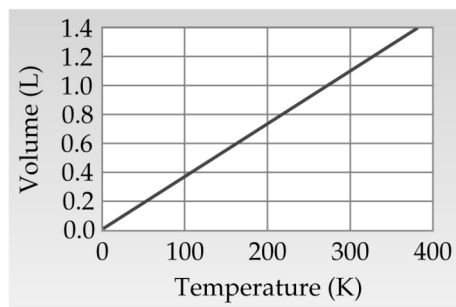
Charles's law: "The volume of a fixed amount of gas at a constant pressure is directly proportional to its Kelvin temperature."

$V \propto T$ (where T must be Kelvin),

$$V_1/T_1 = V_2/T_2$$



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Graphic representation of Charles's law: $V \propto T$, or

$$V_1/T_1 = V_2/T_2$$

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Charles's Law: Example

What will the volume be for 4.0 L of a gas at 25 °C if the temperature is raised to 150 °C?

Must convert temperatures to Kelvin.

$$T_1 = 273 + 25 = 298 \text{ K}$$

$$T_2 = 273 + 150 = 423 \text{ K}$$

Use $V_1 / T_1 = V_2 / T_2$

Charles's Law: Example

Solve for V_2

$$V_2 = \frac{T_2 V_1}{T_1}$$

Substitute in the values

$$V_2 = \frac{(423 \text{ K}) (4.0 \text{ L})}{298 \text{ K}}$$

$$V_2 = 5.7 \text{ L}$$

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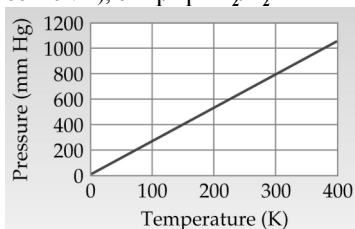
Gay-Lussac's Law: The Relation Between Pressure and Temperature

"The pressure of a gas is directly proportional to its Kelvin temperature for a fixed amount of gas at a constant volume."

$P \propto T$ (where T must be Kelvin), or $P_1/T_1 = P_2/T_2$

Gay-Lussac's law:
 $P \propto T$, or

$P_1/T_1 = P_2/T_2$
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Gay-Lussac's Law: Example

Find the final pressure of a gas at 1.0 atm and 273 K which is raised to 546 K?

Use $P_1/T_1 = P_2/T_2$

Given:

$$T_1 = 273 \text{ K}$$

$$T_2 = 546 \text{ K}$$

$$P_1 = 1.0 \text{ atm}$$

Gay-Lussac's Law: Example

Solve for P_2

$$P_2 = \frac{T_2 P_1}{T_1}$$

Substitute in the values

$$P_2 = \frac{(546 \cancel{\text{K}}) (1.0 \text{ atm})}{273 \cancel{\text{K}}}$$

$$P_2 = 2.0 \text{ atm}$$

The Combined Gas Law

Relationships between P , V and T can be merged into a combined gas law for a fixed amount of gas.

Combined gas law: $P_1 V_1 / T_1 = P_2 V_2 / T_2$

If any five of the six quantities in this equation are known, the sixth can be calculated.

Fixed amount (moles) of gas required

Kelvin temperatures must be used

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Combined Gas Law: Example

A small bubble rises from the bottom of a lake, where the temperature and pressure are 8 °C and 6.4 atm, to the water's surface, where the temperature is 25 °C and pressure is 1.0 atm. Calculate the final volume of the bubble if its initial volume was 2.1 mL.

Convert temperatures to Kelvin!

$$T_1 = 273 + 8 = 281\text{K} \quad T_2 = 273 + 25 = 298\text{K}$$

$$P_1 = 6.4\text{ atm} \quad P_2 = 1.0\text{ atm}$$

$$V_1 = 2.1\text{ mL} = 0.0021\text{ L} \quad V_2 = ?$$

Use $P_1V_1 / T_1 = P_2V_2 / T_2$

Combined Gas Law: Example

Solve for V_2

$$V_2 = \frac{P_1V_1T_2}{T_1P_2}$$

Substitute in the values

$$V_2 = \frac{(6.4\text{ atm})(0.0021\text{ L})(298\text{ K})}{(281\text{ K})(1.0\text{ atm})}$$

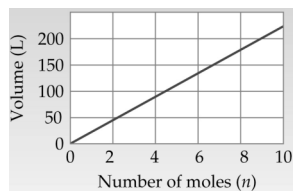
$$V_2 = 0.014\text{ L} = 14\text{ mL}$$

Avogadro's law: The Relation Between Volume and Moles

Avogadro's law: "The volume of a gas is directly proportional to its molar amount at a constant pressure and temperature."

$$V \propto n \text{ (number of moles), or } V_1/n_1 = V_2/n_2$$

Avogadro's law: $V \propto n$,
or $V_1/n_1 = V_2/n_2$

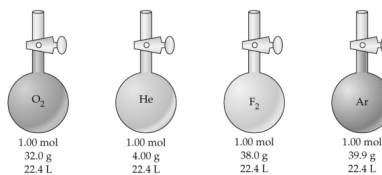


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To simplify comparisons of gas samples, a set of conditions called **standard temperature and pressure (STP)** is defined as follows:

$$\text{STP} = 0^\circ\text{C} (273.15\text{ K}) \text{ and } 1\text{ atm} (760\text{ mm Hg})$$

Standard Molar Volume of any Gas
at STP = 22.4 L/mol



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The Ideal Gas Law

The *Ideal Gas Law* is the relationship that combines the four variables P, V, T, and n for gases:

$$PV/nT = R$$

or

$$PV = nRT$$

R is the **universal gas constant** = 0.0821

If the values of three of the four variables in the ideal gas law are known the fourth can be calculated.

Finding R

The Universal Gas Constant

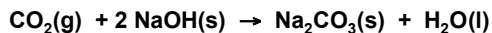
We can easily find what R is by using the fact that 1 mole of gas at STP occupies 22.4 L.

$$R = \frac{PV}{nT} = \frac{(1 \text{ atm})(22.4 \text{ L})}{(1 \text{ mol})(273 \text{ K})} = 0.0821 \text{ L} \cdot \text{atm} / \text{K} \cdot \text{mol}$$

This is a constant for any ideal gas in any conditions.

Example: Ideal Gas Law

A 0.95 L container is filled with CO₂ at 298 K. Solid NaOH is introduced in the container, and 0.88 atm of CO₂ is consumed by the following reaction:

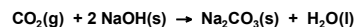


Calculate how many grams of Na₂CO₃ are formed.

Use **PV = nRT** (R = 0.0821)

Example: Ideal Gas Law

Solve for n



$$n = \frac{PV}{RT}$$

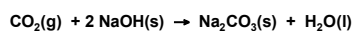
Substitute in the values

$$n = \frac{(0.88 \text{ atm})(0.95 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm} / \text{K} \cdot \text{mol})(298 \text{ K})}$$

$$n = 0.034 \text{ mol CO}_2$$

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Example: Ideal Gas Law



Now using the balanced chemical equation, we see that 1 mol of CO_2 is consumed for every 1 mol of Na_2CO_3 produced.

Thus, the number of moles of Na_2CO_3 is equal to 0.034 moles.

Therefore, using the molecular weight of Na_2CO_3 the total mass of Na_2CO_3 formed in the reaction is:

$$0.034 \text{ mol Na}_2\text{CO}_3 \times \frac{106.0 \text{ g Na}_2\text{CO}_3}{1 \text{ mol Na}_2\text{CO}_3} = 3.6 \text{ g Na}_2\text{CO}_3$$

Test Yourself

How many grams of N_2 are required to fill 2.70×10^4 L to 0.980 atm at 25°C ?

Ideal Gas Law and Density

The *Ideal Gas Law* can be re-written to express the density of a gas:

$$PM = dRT$$

where: d = density of gas (in g/L)

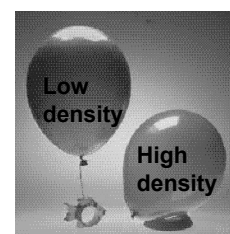
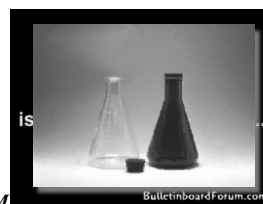
M = molar mass of gas (in g/mol)

Note that as molar mass increases, density increases.

Example: Neon (20.18 g/mol) is a denser gas than helium (4.003 g/mol)

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GAS DENSITY



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USING GAS DENSITY

The density of air at 15 °C and 1.00 atm is 1.23 g/L. What is the molar mass of air?

Answer: Calculate the molar mass

$$d = 1.23 \text{ g/L} \quad P = 1.00 \text{ atm} \quad T = 288 \text{ K}$$

$$M = dRT/P$$

$$M = (1.23 \text{ g/L})(0.0821)(288 \text{ K})/(1.00 \text{ atm})$$

$$M = dRT/P = 29.1 \text{ g/mol}$$

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Partial Pressure and Dalton's Law

In a gas, molecules are so far apart that they have little or no effect on each other. In a *mixture* of gases, each molecule acts independently of all the others.

Dalton saw that the collisions of each molecule against the walls of its container contributed to the *total* pressure in the container.

Dalton's Law of Partial Pressure states that the total pressure (P_{total}) of a mixture of gases is the sum of the individual partial pressures. It is written:

$$P_{\text{total}} = P_{\text{gas1}} + P_{\text{gas2}} + P_{\text{gas3}} + \dots$$

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Example: Partial Pressure

To a tank containing N_2 at 2.0 atm and O_2 at 1.5 atm we add an unknown quantity of CO_2 until the total pressure is 4.3 atm. What is the partial pressure of the CO_2 ?

Using Dalton's law:

$$P_{\text{T}} = P_{\text{O}_2} + P_{\text{N}_2} + P_{\text{CO}_2}$$

Now plugging in the given values:

$$4.3 = 1.5 + 2.0 + P_{\text{CO}_2}$$

Solving for P_{CO_2} , you find that the partial pressure of CO_2 is 0.8 atm.

Intermolecular Forces

The biggest difference between the gaseous state and the condensed states is the effect of **intermolecular forces**.

Intermolecular forces are *weak electrostatic forces* of varying strength between molecules.

Three major types of intermolecular forces:

- Dipole-Dipole
- London Dispersion
- Hydrogen Bonding

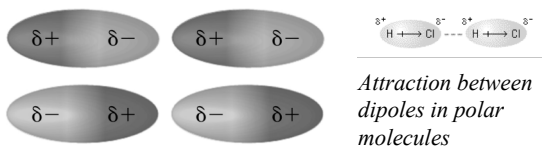
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Dipole-Dipole forces are moderately strong electrostatic attractions between molecules with permanent dipoles.

Molecules tend to line up according to dipole since the positive end of one molecule attracts the negative end of another.

They can exist between two identical molecules *or* two different molecules.

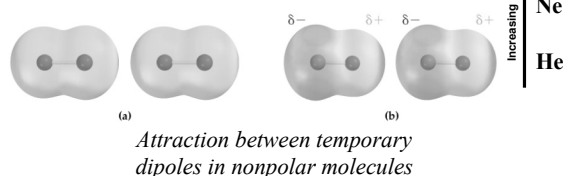


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London dispersion forces are very weak *temporary* electrostatic attraction between molecules.

ALL molecules have London Dispersion forces; however, London Dispersion forces are the *weakest* intermolecular force.

The London Dispersion Force increases in strength as the size of the molecule increases.



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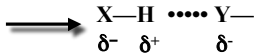
Hydrogen bonds: A hydrogen bond is a special form of dipole-dipole which creates very strong attractions between H atoms bonded to N, O, or F *and* the electrons of N, O, or F on adjacent molecules.

Two important conditions must be met for hydrogen bonding to occur:

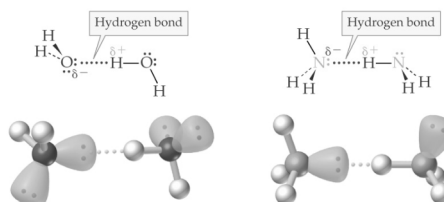
- * One molecule has a hydrogen atom attached by a covalent bond to an atom of N, O, or F.

- * The other molecule has an N, O, or F atom.

Hydrogen bonding only when X and Y are either a N, O or F

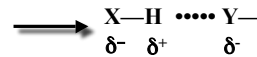


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Hydrogen bonding in water and ammonia

Hydrogen bonding only when X and Y are either a N, O or F



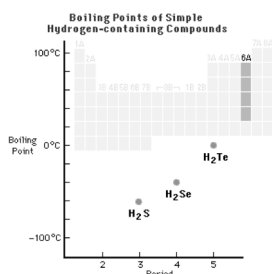
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Hydrogen bonds can be quite strong.

Since H_2O molecules held together by H bonds, more energy must be added to make them boil (break IM bonds).

Hence, boiling point of H_2O is much higher than other related Group 6A compounds.



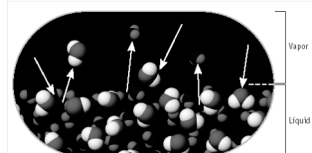
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Liquids

Molecules are in constant motion as liquids.

If a molecule is near the surface of a liquid *and* if it has enough energy, it can break free of the liquid and escape into the gas state called **vapor**.

In open containers, the gaseous molecule will wander away from the liquid, and this process will continue until all the molecules escape from the container. This is what happens during **evaporation**.

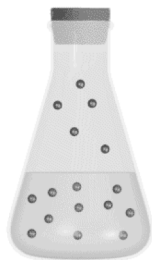


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In a *closed container*, gaseous molecules cannot escape. Random motions of molecule occasionally will bring them gaseous molecules back into the liquid - **condensation**.

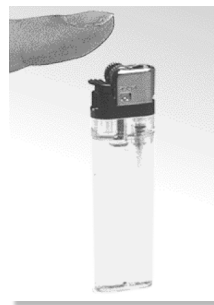
After awhile, the number of molecules reentering the liquid from the gas becomes equal to the number of molecules escaping the liquid into the gas.

This is called **dynamic equilibrium**. Evaporation and condensation take place at the same rate as long as temperature remains constant.



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Dynamic Equilibrium



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In closed containers at equilibrium, gas molecules contribute to the total pressure of the gas above the liquid by Dalton's law.

This is the **vapor pressure** (VP) of the liquid.

Vapor pressure depends on temperature and the intermolecular forces of compound.

Vapor pressure rises with increasing temperature and ultimately becomes equal to the atmospheric pressure.

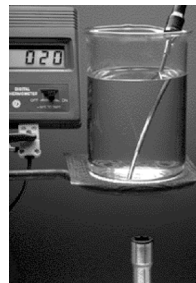
When VP = atmospheric pressure (760 mm Hg), **boiling** occurs, and this temperature is called the **normal boiling point**.

Boiling Points depend on atmospheric pressure!

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If atmospheric pressure is high (Death Valley, etc.), boiling point is high.

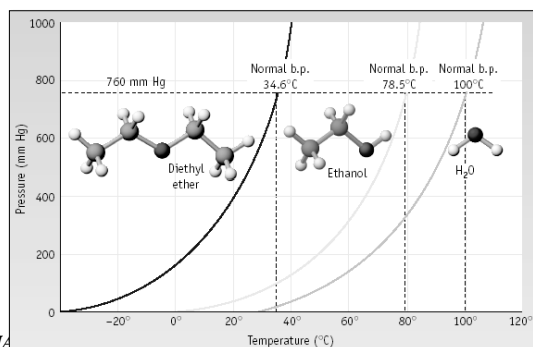
If atmospheric pressure is low (Colorado, etc.), boiling point is low.



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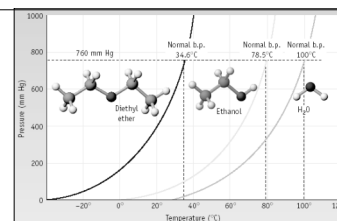


Vapor Pressure



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Vapor Pressure (VP)



The curves (VP vs T) show the VP for different liquids at various T

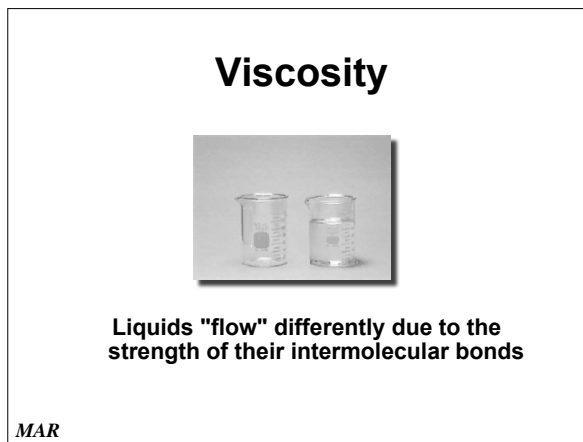
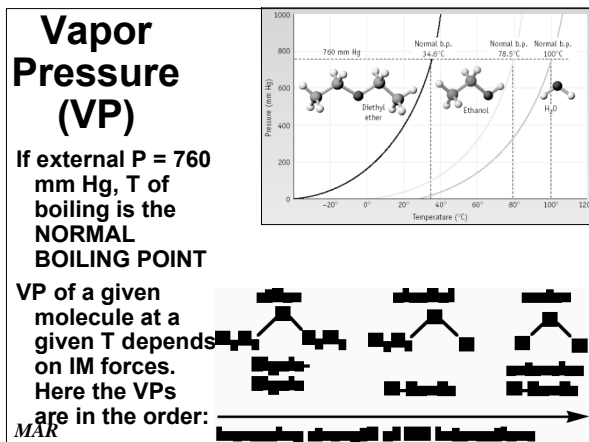
The VP rises with T.

When VP = external P, the liquid boils.

This means that BPs of liquids change with altitude.

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- ## IM Forces / VP Summary
- **Intermolecular (IM) Forces** turn gases (which have no IM forces) into liquids and solids (which have lots of IM forces)
 - Stronger IM forces lead to greater viscosity, larger surface tension, higher boiling points, lower volatility
 - **Vapor Pressure (VP)** depends on IM forces. As Temperature increases, VP increases
 - The higher the IM force, the lower the VP (at a given Temperature)
 - When VP equals external pressure, boiling occurs. The temperature at which a liquid boils at 1 atmosphere of pressure is called the **normal boiling point** of the liquid
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Water: A Unique Liquid

Water in our blood transports substances throughout our body. All biochemical reactions in our body are carried out in water.

Water has a high **specific heat capacity** (absorbs large quantities of heat while changing its temperature very little).

Water has unusually high **heat of vaporization** – it carries away large amount of heat when it evaporates.

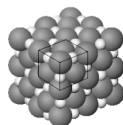


Solids

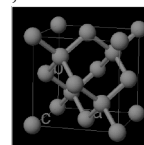
Solids exist in two forms: amorphous solids and crystalline solids.

Amorphous solids have randomly scattered particles (atoms, ions or molecules) with no long range structure. *Examples:* glass, tar, etc.

Crystalline solids have particles (atoms, ions or molecules) with an ordered arrangement extending over a long range. *Examples:* ice, NaCl, diamond, etc.



NaCl



Diamond

Crystalline solids can be categorized as:

- **Ionic** such as $\text{Na}^+ \text{Cl}^-$ whose constituent particles are ions.
- **Covalent** such as *diamond* or *quartz* where molecules are held together by covalent bonds.
- **Molecular** such as *sucrose* or *ice* whose constituent particles are molecules held together by the intermolecular forces.
- **Metallic** such as *silver* or *iron* with a three dimensional array of metal ions immersed in electrons that are free to move about.
- More on solids in "**Solids**" lab

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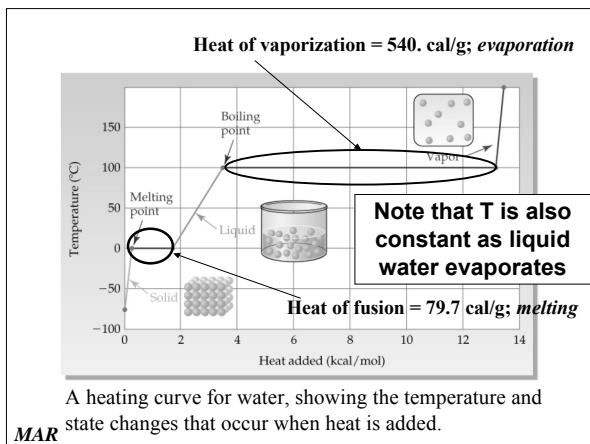
Changes of States

Heat of fusion: The quantity of heat required to completely melt a substance once it has reached its melting point temperature.

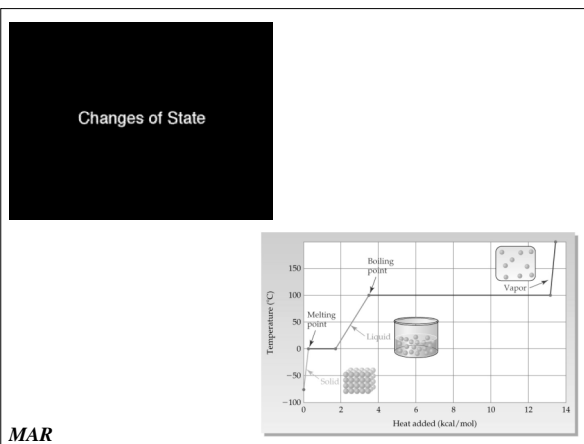
Heat of vaporization: The quantity of heat required to completely vaporize a substance once it has reached its boiling point temperature.

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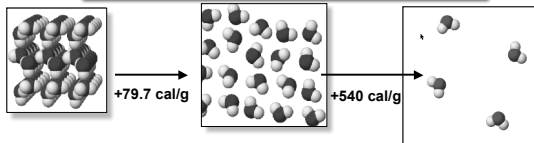


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Heat & Changes of State

What quantity of heat is required to melt 500. g of ice and heat the water to steam at 100. °C?

Heat of fusion of ice = 79.7 cal/g
Specific heat of water = 1.00 cal/g·°C
Heat of vaporization = 540. cal/g



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Heat & Changes of State

What quantity of heat is required to melt 500. g of ice and heat the water to steam at 100. °C?

1. To melt ice:

$$q = (500. \text{ g})(79.7 \text{ cal/g}) = 3.99 \times 10^4 \text{ cal}$$

2. To raise water from 0 °C to 100. °C:

$$q = 500. \text{ g} \cdot 1.00 \text{ cal/g} \cdot ^\circ\text{C} \cdot (100. - 0)^\circ\text{C} = 5.00 \times 10^4 \text{ cal}$$

3. To evaporate water at 100. °C:

$$q = (500. \text{ g})(540. \text{ cal/g}) = 2.70 \times 10^5 \text{ cal}$$

4. Total heat energy = $3.60 \times 10^5 \text{ cal} = 360. \text{ kcal}$

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End of Chapter 7

To review and study for Chapter 7, look at the "Concepts to Remember" at the end of Chapter Seven