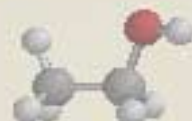
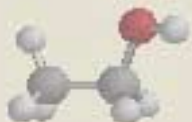
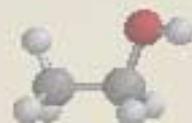
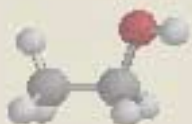


Chem101 - Lecture 6

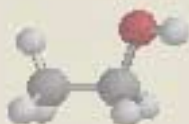
States of Matter





Introduction

- Matter has three possible states:
 - Solid
 - Liquid
 - Gas
- We will investigate the differences in the physical properties exhibited by each of these states
 - We will work to gain understanding of the forces and interactions between molecules that determine these states.

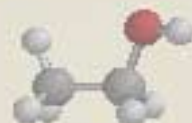
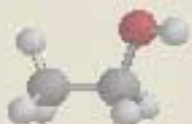


Introduction

- In Chapter 4 (Section 4.11) we saw that the state pure substances depends on the temperature.

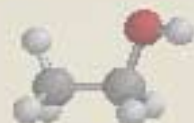
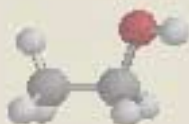
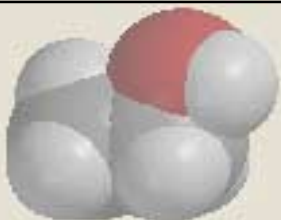
■ TABLE 4.8 Behavior of selected pure substances on heating

Temperature (°C)	Behavior or state of substance					
	Oxygen (O ₂)	Carbon monoxide (CO)	Water (H ₂ O)	Salt (NaCl)	Copper (Cu)	Quartz (SiO ₂)
-220	Solid	Solid	Solid	Solid	Solid	Solid
-218	Melts	Solid	Solid	Solid	Solid	Solid
-199	Liquid	Melts	Solid	Solid	Solid	Solid
-192	Liquid	Boils	Solid	Solid	Solid	Solid
-183	Boils	Gas	Solid	Solid	Solid	Solid
0	Gas	Gas	Melts	Solid	Solid	Solid
100	Gas	Gas	Boils	Solid	Solid	Solid
801	Gas	Gas	Gas	Melts	Solid	Solid
1083	Gas	Gas	Gas	Liquid	Melts	Solid
1413	Gas	Gas	Gas	Boils	Liquid	Solid
1610	Gas	Gas	Gas	Gas	Liquid	Melts
2230	Gas	Gas	Gas	Gas	Liquid	Boils
2595	Gas	Gas	Gas	Gas	Boils	Gas
2600	Gas	Gas	Gas	Gas	Gas	Gas



Introduction

- We will also see that it depends on pressure.
- Interparticle forces were also discussed in Chapter 4 (Section 4.11).
 - We will see that the magnitude of these forces influences the state of matter.



Introduction

- At normal temperatures and pressures, the pure forms of most of the elements are solids.

- There are only two liquids

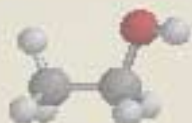
Bromine (Br) and mercury (Hg).

- And eleven gases

The noble gases, helium(He), neon (Ne), argon(Ar), krypton(Kr), xenon(Xe), radon(Rn)

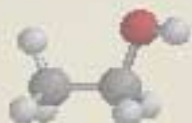
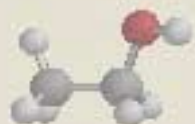
Plus a group of elements located in the upper right-hand corner of the periodic table:

hydrogen(H_2), fluorine(F_2), chlorine(Cl_2), oxygen(O_2) and Nitrogen(N_2)



Observed Properties of Matter

- The states of matter can be readily distinguished from one another based on four properties:
 - **Density**
 - ⌘ Density is a characteristic property of pure substances
 - ⌘ Solids and liquids are much more dense than gases.
 - Solid and liquid have very similar densities.
 - The solid is usually slightly more dense than the liquid.
 - Water is an exception, which is why ice floats.

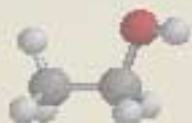
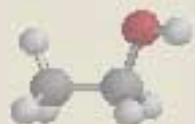


Observed Properties of Matter

- The states of matter can be readily distinguished from one another based on four properties:

- **Shape**

- ⌘ Solids have well defined shapes and require no container to hold them.
 - ⌘ Liquids flow and take on the shape of their container
 - The volume of a container that is occupied by a liquid depends on the amount of liquid placed in the container.



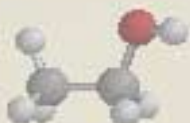
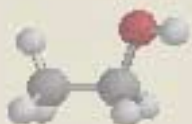
Observed Properties of Matter

- The states of matter can be readily distinguished from one another based on four properties:

- **Shape**

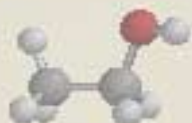
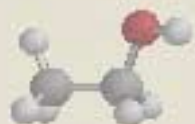
- ⌘ Gases , like liquids, flow and take on the shape of their container

- Unlike liquids, gases expand to occupy the total volume of a container, regardless of the quantity of gas placed in the container.



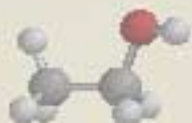
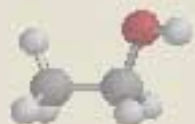
Observed Properties of Matter

- The states of matter can be readily distinguished from one another based on four properties:
 - **Compressibility** – is the ability to reduce the volume of a substance by applying pressure
 - Gases are highly compressible
 - Liquids and solids both have very low compressibilities.



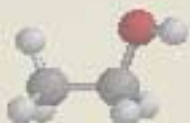
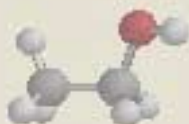
Observed Properties of Matter

- The states of matter can be readily distinguished from one another based on four properties:
 - **Thermal expansion** – is the increase in volume that occurs when a substance is heated.
 - Gases display moderately high thermal expansion
 - Liquids and Solids expand very little when heated.



Kinetic-Molecular Theory of Matter

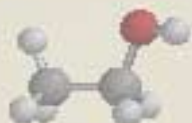
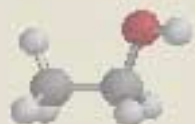
- The **kinetic-molecular theory of matter** is the theory used to explain the different states of matter.
- Ancient Greeks proposed pieces of the theory over 2400 years ago.



Kinetic-Molecular Theory of Matter

1. Matter is composed of tiny particles called **molecules**.
2. The particles possess **kinetic energy** because they are moving.
3. The particles possess **potential energy** because they are attracted to and repelled by each other.
4. The average speed of a particle (kinetic energy) increases with **temperature**.
5. The particles transfer energy from one to another during collisions.

A Kinetic-Molecular Simulation of matter

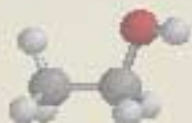
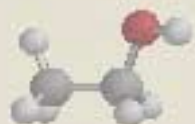


Kinetic-Molecular Theory of Matter

- The total energy of a sample of matter is equal to the sum of its kinetic and potential energies.
 - **Kinetic energy** (K.E.) is the energy a particle possesses because it is moving:

$$K.E. = \frac{1}{2}mv^2$$

m is the mass of the particle, and v is its velocity.

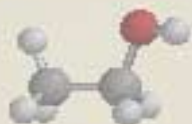
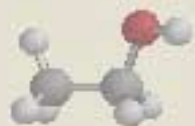


Kinetic-Molecular Theory of Matter

- The total energy of a sample of matter is equal to the sum of its kinetic and potential energies.
 - **Potential energy** is the energy a particle possesses due to its position relative to the other particles.

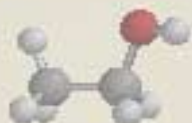
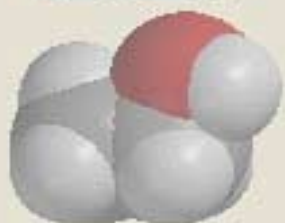
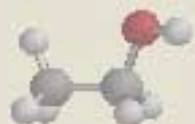
The potential energy from attractive interactions increases as the particles become further apart from one another.

The potential energy from repulsive interactions increases as the particles become closer together.

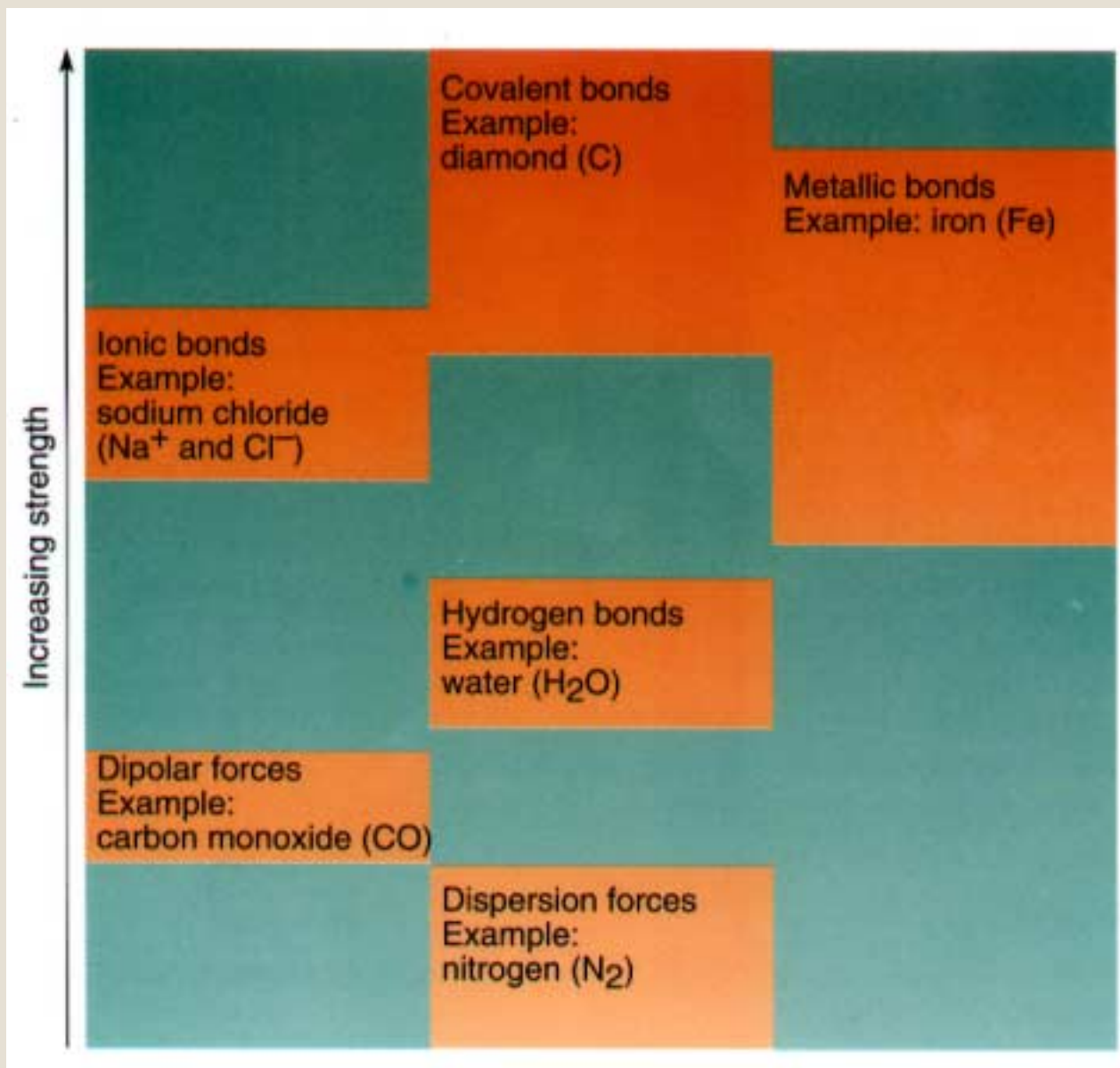


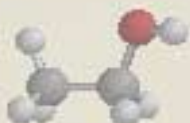
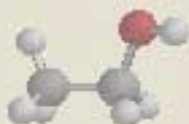
Kinetic-Molecular Theory of Matter

- Molecules experience an array of repulsive and attractive interactions
 - These were discussed in Chapter 4 (Section 4.11).



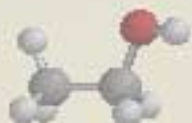
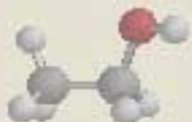
Kinetic-Molecular Theory of Matter





Kinetic-Molecular Theory of Matter

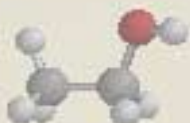
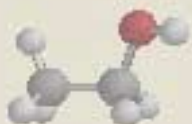
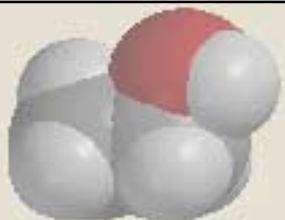
- The states of matter can be understood by considering the **cohesive** and **disruptive** influences that kinetic and potential energies have on a collection of particles:
 - Cohesive forces hold particles together;
They arise from attractive interactions (potential energy).
 - Disruptive forces pull particles apart
They arise from kinetic energy and counter the effects of the cohesive forces.



Kinetic-Molecular Theory of Matter

- Disruptive forces are directly dependent on temperature
- Cohesive interactions are independent of temperature.
 - As the temperature increases the disruptive forces become greater and more dominant.

A Kinetic-Molecular Simulation of matter



Kinetic-Molecular Theory of Matter

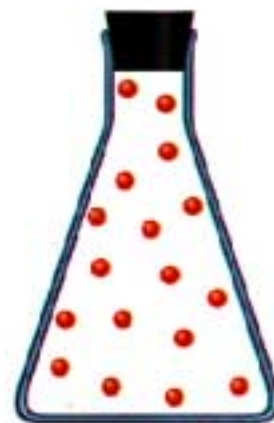
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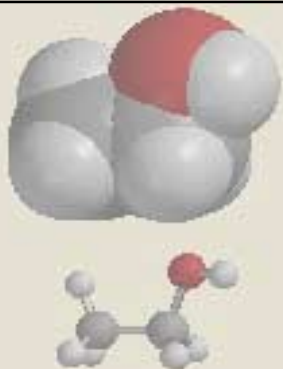
(a) Solid state: The particles are close together and held in fixed positions; they do not need a container.



(b) Liquid state: The particles are close together but not held in fixed positions; they take shape of the container.

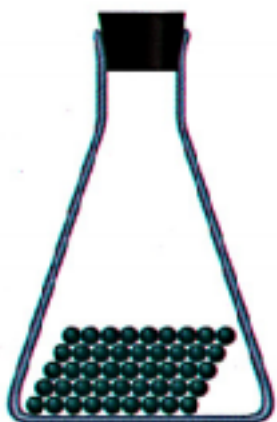


(c) Gaseous state: The particles are far apart and completely fill the container.



Solid State

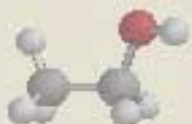
- In solids the cohesive forces are much greater than the disruptive forces.
 - In a crystalline solid, each molecule is held at a fixed location within a crystal lattice.
 - The disruptive forces cause the molecules to vibrate at their fixed location, but does not allow the molecules to move past one another.



(a) Solid state: The particles are close together and held in fixed positions; they do not need a container.

A Kinetic-Molecular Simulation of matter





Solid State

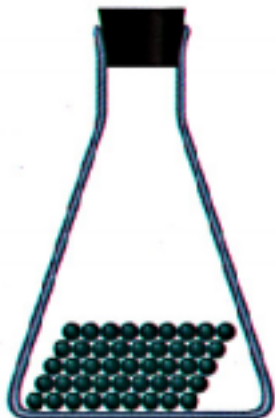
- The kinetic-molecular model of matter can explain the characteristic properties displayed by solids.

- *High density*

Because the cohesive forces are dominant, the molecules in a solid are very close together, which results in a greater number of molecules in a given volume.

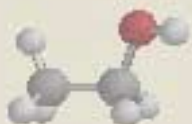
- *Definite shape*

Because the molecules in a solid cannot move past one another, they maintain a defined shape.



(a) Solid state: The particles are close together and held in fixed positions; they do not need a container.





Solid State

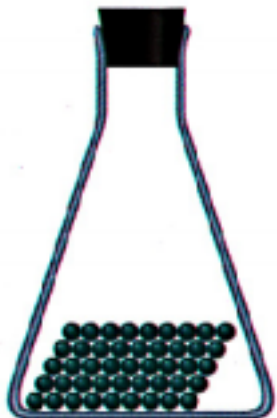
- The kinetic-molecular model of matter can explain the characteristic properties displayed by solids.

- *Low compressibility*

Because the molecules in a solid are very close together, increasing the pressure is unable to move them much closer together.

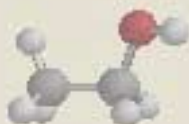
- *Small thermal expansion*

Though heating a solid causes the molecules to vibrate more about their fixed positions, the cohesive force still predominates, holding them very close together.



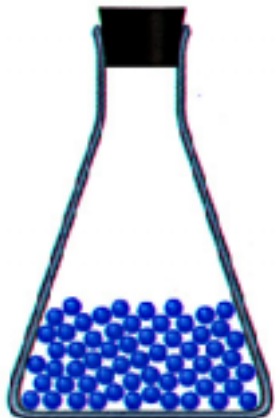
(a) Solid state: The particles are close together and held in fixed positions; they do not need a container.





Liquid State

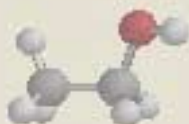
- In the liquid state the cohesive forces still predominate, holding the molecules almost as close together as in the solid state.
 - The disruptive forces are strong enough, though, that the molecules do not have fixed locations.
 - Instead they are able to move about, slipping past one another.



(b) Liquid state: The particles are close together but not held in fixed positions; they take shape of the container.

A Kinetic-Molecular Simulation of matter





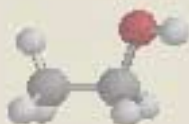
Liquid State

- Because the cohesive forces are able to still hold the molecules close together in liquid, their densities, compressibilities and thermal expansions are similar to their corresponding solids
 - All three of these properties depend on the distances between molecules.
- Unlike solids, liquids do not have defined shapes
 - The molecules are capable of sliding past one another.



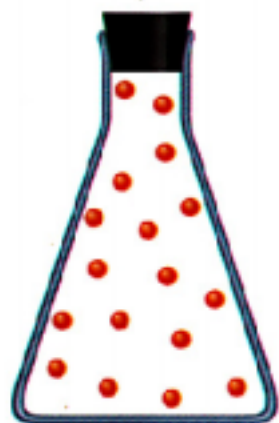
(b) Liquid state: The particles are close together but not held in fixed positions; they take shape of the container.





Gas State

- In the gas state, the disruptive interactions predominate; the cohesive forces are no longer able to hold the molecules together.
 - In gases the molecules move about independently of one another.
 - They come in contact with one another only when they collide.



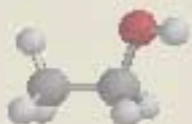
(c) Gaseous state: The particles are far apart and completely fill the container.

And then only briefly before heading off in a new direction.

Between collisions the molecules travel in straight lines at a constant velocity.

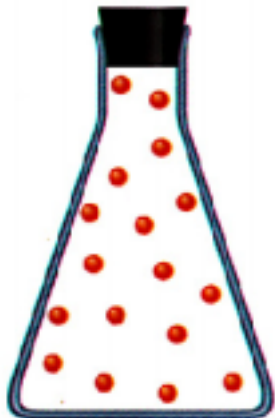
A Kinetic-Molecular Simulation of matter





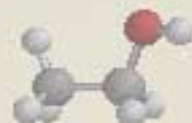
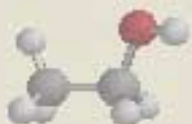
Gas State

- The kinetic-molecular model of matter can explain the characteristic properties displayed by gases.
 - *Low density* – because the disruptive forces are dominant, the molecules are spread out and as far apart from one another as possible. This gives gases very low densities, and therefore little mass in a given volume
 - *Indefinite shape* – Like liquids, the molecules can easily slide past one another. They take on the shape of their container.



(c) Gaseous state: The particles are far apart and completely fill the container.





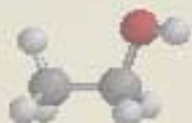
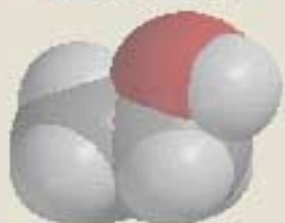
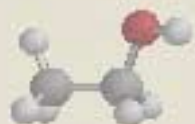
Gas Laws

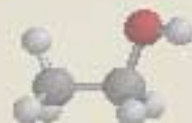
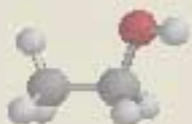
- For gases, simple mathematical equations can be derived to relate a gas's *volume*, *pressure* and *temperature*.
 - These equations are called **state equations** because they describe the state of the gas.
 - Because the molecules in a gas do not interact strongly with one another, the identity of a gas is not important: the equations apply to all gases in the same way.



Gas Laws

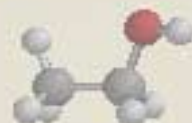
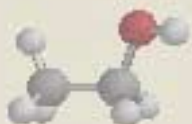
- The same cannot be said for liquids and solids.
 - Because the molecules in liquids and solids do interact strongly with one another, the relationship between pressure, volume and temperature are much more complicated and dependent on the identity of the substance.





Gas Laws

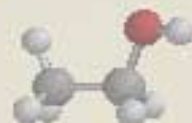
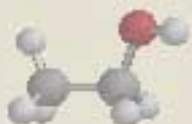
- **Pressure** is defined as the force applied per unit area to a surface.
 - The pressures exerted by gases can be measured with a device called a manometer.
 - A manometer that is used to measure atmospheric pressure is called a **barometer**.



Gas Laws

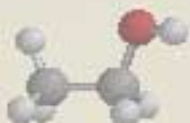
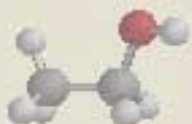
- This figure shows how to make a simple barometer, which was invented by Torricelli in the 1600's.





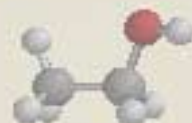
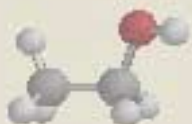
Gas Laws

- Torricelli's barometer measures pressure by determining the height of a column of mercury that the pressure from the atmosphere can support.
 - Using this kind of barometer, the pressures are measured in units of millimeters of mercury (*mm-Hg*) or *Torr* (in honor of Torricelli).



Gas Laws

- There is a pressure that is defined as a standard pressure.
 - It is approximately equal to the average pressure exerted by the earth's atmosphere at sea level and is defined as 1 *atmosphere* (1 *atm*).
 - 1 *atm* is equal to 760 *Torr*.

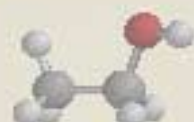
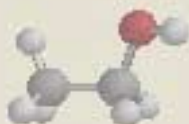


Gas Laws

- The **temperature** scale that must be used in gas law equations is the **Kelvin** scale.
 - The **Kelvin** scale has the same sized degrees as the **Celsius** scale, but it is shifted so that 0 K is at absolute zero.

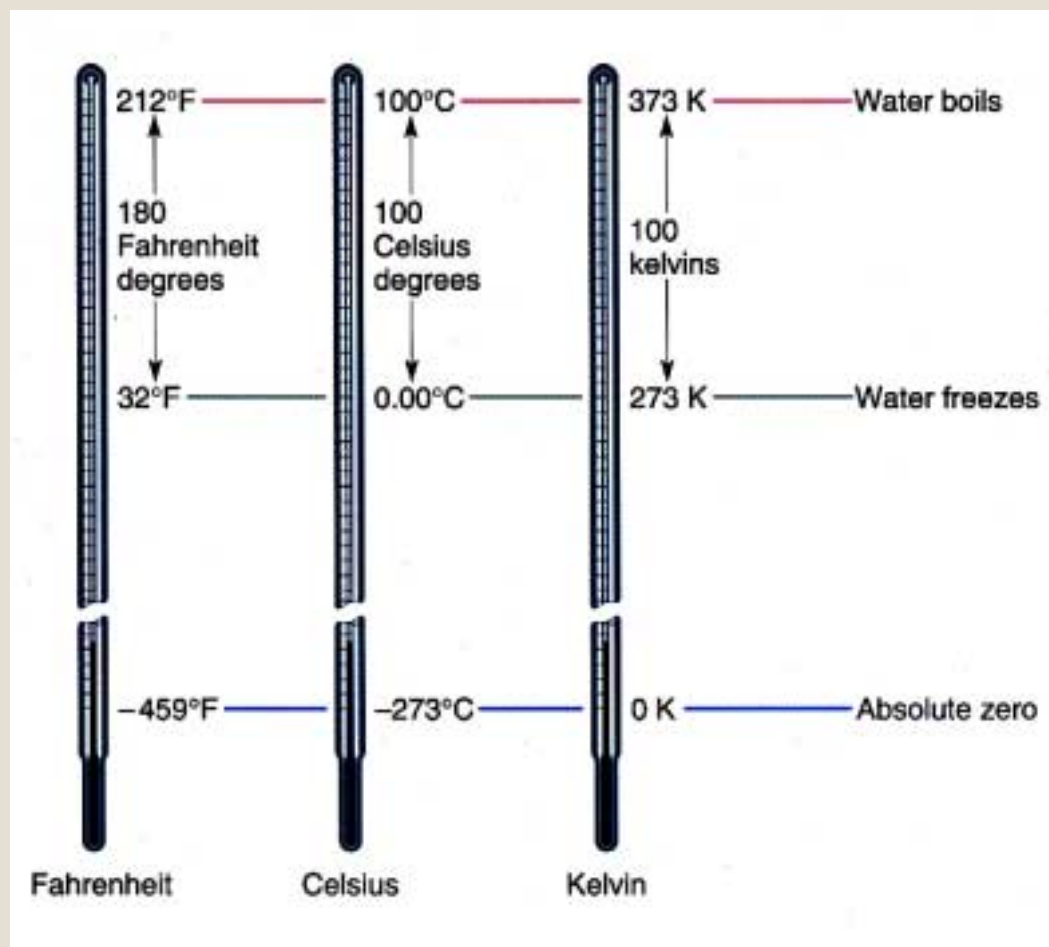
This is the temperature at which the kinetic energy is zero.

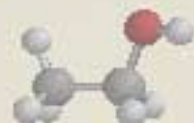
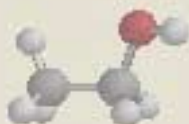
All motion stops.
 - The shift is 273° , that is $0\text{ K} = -273^\circ\text{C}$



Gas Laws

- The **temperature** scale that must be used in gas law equations is the **Kelvin** scale.





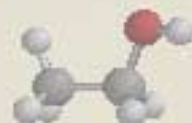
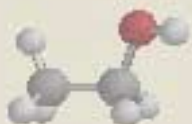
P , T and V Relationships in Gases

- **Boyle's Law**

- In 1662, the Irish chemist Robert Boyle derived an equation that relates a gas's pressure to its volume when the temperature and the number of gas molecules is held constant:

$$P = \frac{k}{V}$$

- Where k is a constant



P , T and V Relationships in Gases

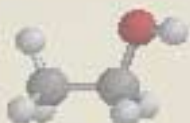
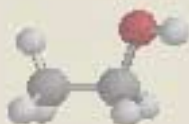
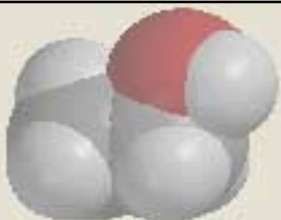
- **Boyle's Law**

- Boyle's law states that the pressure of a gas is inversely proportional to the volume of the gas.

$$P = \frac{k}{V}$$

- A rearrangement of this equation gives:

$$PV = k$$



P , T and V Relationships in Gases

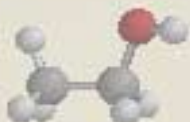
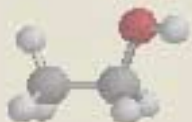
- **Charles's Law**

- In 1787, the French scientist Jacques Charles derived an equation that relates a gas's volume to its temperature when its pressure is held constant.

$$V = k'T$$

- Where k' is a constant.

The prime sign after the k is to distinguish this constant from the one that appeared in Boyle's law equation.

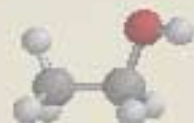
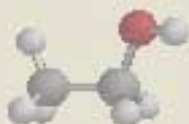


P , T and V Relationships in Gases

- **Charles's Law**

- A rearrangement of this equation gives:

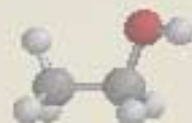
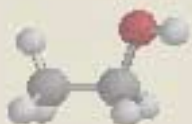
$$\frac{V}{T} = k'$$



P , T and V Relationships in Gases

- **Combined Gas Law**
 - Boyle and Charles' Laws can be combined to give the combined gas law equation:

$$\frac{PV}{T} = k''$$



P , T and V Relationships in Gases

- **Combined Gas Law**

- Initial and final states

If a gas is changed from an initial (i) to a final (f) state, then there will be a set of values for P , V , and T , for each state:

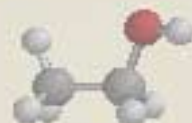
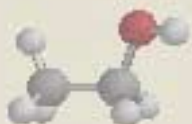
Initial state -. P_i , V_i , T_i

Final state - P_f , V_f , T_f

Applying the combined gas law to each:

Initial state:
$$\frac{P_i V_i}{T_i} = k''$$

Final state:
$$\frac{P_f V_f}{T_f} = k''$$



P , T and V Relationships in Gases

- **Combined Gas Law**

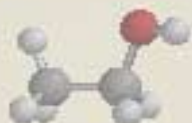
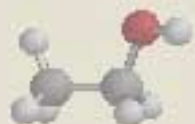
- Since both are equal to k , we can set them equal to each other:

$$\frac{P_i V_i}{T_i} = \frac{P_f V_f}{T_f}$$

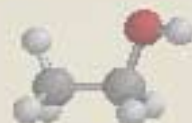
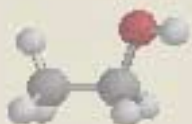
- This equation relates the initial and final states of a gas when any or all of the conditions of pressure, volume or temperature change.



Exercise 6.25

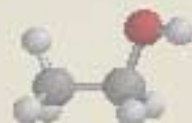
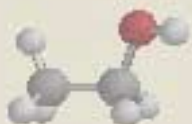


A 200 *mL* sample of oxygen gas is collected at 26°C and a pressure of 690 Torr. What volume will the gas occupy at STP (0°C and 760 Torr)?



Ideal Gas Law

- The combined gas law works only if the number of gas molecules remains unchanged.
- **Avogadro's Law**
 - *Equal volumes of different gases measured at the same temperature and pressure contain equal numbers of molecules of gas.*

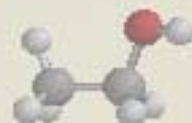
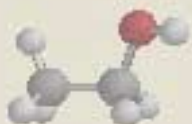


Ideal Gas Law

- **Avogadro's Law**

- This means that the constant in the combined gas law equation, k'' , is directly proportional to the number of gas molecules:

$$k'' = nk'''$$



Ideal Gas Law

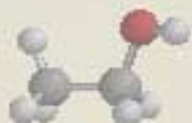
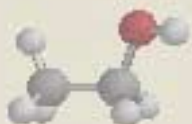
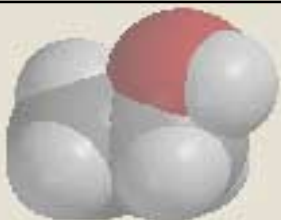
- Combination of Boyle's, Charles's and Avogadro's Laws give the Ideal gas law equation:

$$\frac{PV}{T} = k''$$

$$\frac{PV}{T} = nk'''$$

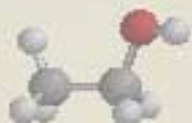
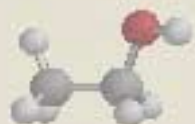
$$\frac{PV}{T} = nR \text{ (Ideal Gas Law Equation)}$$

$$PV = nRT$$



Ideal Gas Law

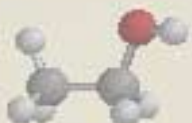
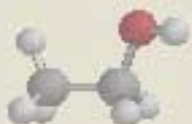
- $R = k'''$ is the *Ideal Gas Law Constant* or *Universal Gas Constant*
 - $R = 0.0821 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$
 - RT is a measure of the kinetic energy per mole of gas
- An ideal gas has no potential energy
 - It is neither attracted to the itself or to the walls of the container



Dalton's Law

- For ideal gases (no potential energy) at a fixed volume and temperature, the pressure exerted by the gas on the walls of the container is proportional to number of moles of gas present:

$$P = \frac{nRT}{V}$$

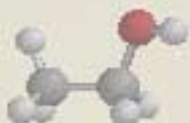
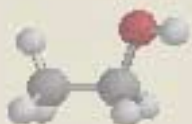


Dalton's Law

- Dalton's Law of partial pressures states that for a mixture of gases at a fixed volume and temperature, each gas will contribute in proportion to its number to the total pressure:

$$P_{tot} = \frac{n_A RT}{V} + \frac{n_B RT}{V} + \frac{n_C RT}{V} + \dots$$

$$P_{tot} = P_A + P_B + P_C + \dots$$



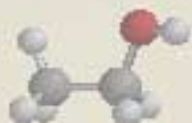
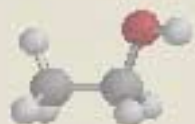
Dalton's Law

- The contribution made by each gas is called its partial pressure

$$P_{tot} = \frac{n_A RT}{V} + \frac{n_B RT}{V} + \frac{n_C RT}{V} + \dots$$

$$P_{tot} = P_A + P_B + P_C + \dots$$

- In lab, you made use of this relationship in the gas laws experiment when you subtracted the partial pressure of water vapor from the total pressure to get the partial pressure of the CO₂.

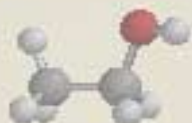
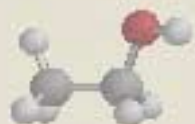


Graham's Law

- **Effusion** is the escape of a gas through a small hole in a container.
- **Diffusion** is the spontaneous mixing of gases when brought together.
 - At a fixed temperature, the average kinetic energy of all gases is the same, regardless of its identity:
 - Graham's Law states that the effusion rate of a gas is inversely proportional to the square root of its molecular mass



Graham's Law



$$(K.E.)_A = \frac{1}{2} m_A v_A^2$$

$$(K.E.)_B = \frac{1}{2} m_B v_B^2$$

$$(K.E.)_A = (K.E.)_B \quad (\text{when } T_A = T_B)$$

$$\frac{1}{2} m_A v_A^2 = \frac{1}{2} m_B v_B^2$$

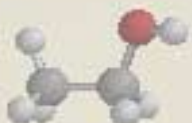
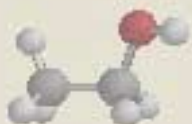
$$m_A v_A^2 = m_B v_B^2$$

$$\frac{v_A^2}{v_B^2} = \frac{m_B}{m_A}$$

$$\left(\frac{v_A}{v_B} \right)^2 = \frac{m_B}{m_A}$$

$$\frac{v_A}{v_B} = \sqrt{\frac{m_B}{m_A}}$$

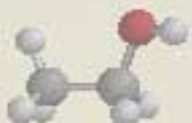
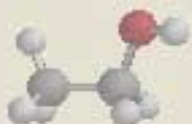
$$\frac{\text{effusion rate } A}{\text{effusion rate } B} = \sqrt{\frac{\text{molecular mass } B}{\text{molecular mass } A}} \quad (\text{Graham's Law})$$



Graham's Law

- Graham's Law states that the effusion rate of a gas is inversely proportional to the square root of its molecular mass.

$$\frac{\text{effusion rate } A}{\text{effusion rate } B} = \sqrt{\frac{\text{molecular mass } B}{\text{molecular mass } A}} \quad (\text{Graham's Law})$$

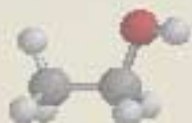
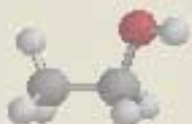


Changes in State

- The state of matter is determined by the competition between potential energy (cohesive interactions) and kinetic energy (disruptive interactions).
 - Kinetic energy is proportional to temperature and independent of the identity of the molecules

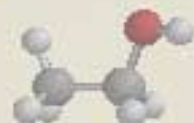
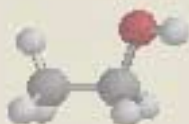
The higher the temperature, the greater the kinetic energy, and the greater the molecule's velocity

This disrupts the cohesive (potential energy) interactions.



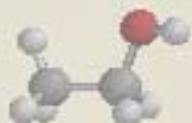
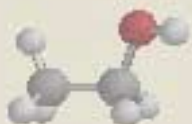
Changes in State

- Kinetic energy is proportional to temperature and independent of the identity of the molecules
 - The higher the temperature, the greater the kinetic energy, and the greater the molecule's velocity
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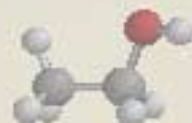
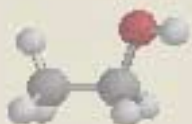
Changes in State

- Potential energy is not influenced by the temperature.
 - The contributions to the potential energy include:
 - Covalent bonds
 - Metallic bonds
 - Ionic bonds
 - Hydrogen bonds
 - Dipole interactions
 - Dispersion interactions.
 - These are arranged in descending order of strength.



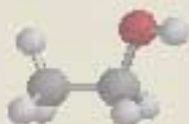
Changes in State

- The stronger the cohesive energy holding the molecules together
 - the higher the temperature (greater the kinetic energy) required to disrupt these interactions
 - hence, the greater the melting point and greater the boiling point.



Evaporation and Vapor Pressure

- Evaporation or vaporization
 - Molecules from a liquid or solid continually escape the surface to become gas molecules.
- Condensation
 - The opposite process where gas molecules strike and stick to the surface of the liquid.



Evaporation and Vapor Pressure

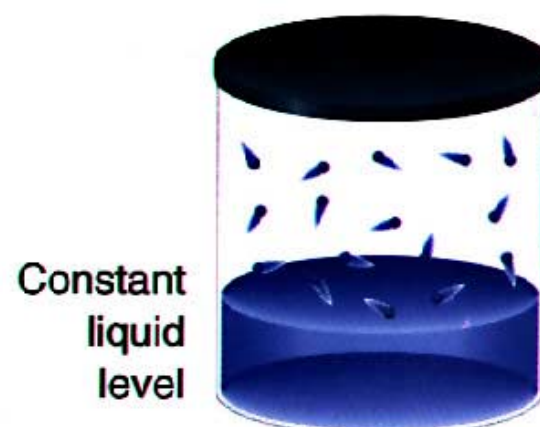
- At equilibrium the rate of evaporation and condensation are equal.
 - This occurs when a lid is placed on the container containing the liquid or gas so that the gas molecules cannot escape.



Initially

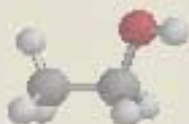


After some time



Constant
liquid
level

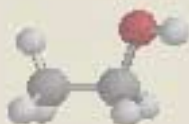
At equilibrium



Evaporation and Vapor Pressure

- Vapor pressure
 - The vapor pressure is the pressure of the vapor at equilibrium.
 - Different liquids have different vapor pressures.

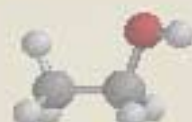
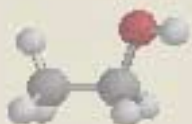
Liquid	Molecular weight (u)	Polarity	Vapor pressure (torr)
pentane (C ₅ H ₁₂)	72	Nonpolar	414.5
hexane (C ₆ H ₁₄)	86	Nonpolar	113.9
heptane (C ₇ H ₁₆)	100	Nonpolar	37.2
ethanol (C ₂ H ₅ — OH)	46	Polar (hydrogen bonds)	43.9
1-propanol (C ₃ H ₇ — OH)	60	Polar (hydrogen bonds)	17.3
1-butanol (C ₄ H ₉ — OH)	74	Polar (hydrogen bonds)	7.1



Evaporation and Vapor Pressure

- Vapor pressure
 - The vapor pressures depend on the strength of the cohesive intermolecular interactions which hold the molecules together in the liquid state.

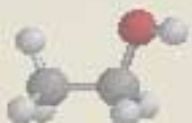
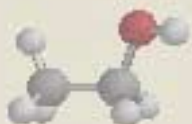
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Evaporation and Vapor Pressure

- The vapor pressure depends on temperature.
 - The higher the temperature the higher the vapor pressure

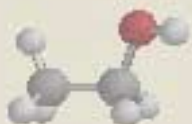
Temperature (°C)	Vapor pressure (torr)
0	4.6
20	17.5
40	55.3
60	149.2
80	355.5
100	760.0



Boiling and Boiling Points

- A liquid boils when its vapor pressure equals the atmospheric pressure.
 - For example, at sea level water boils at 100°C because its vapor pressure is equal to 760 *Torr* (= 1 *atm*) at that temperature

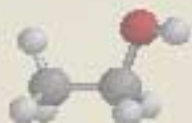
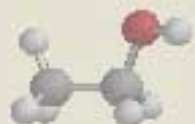
Temperature (°C)	Vapor pressure (torr)
0	4.6
20	17.5
40	55.3
60	149.2
80	355.5
100	760.0



Boiling and Boiling Points

- Water boils at lower temperatures at higher elevations because the atmospheric pressure drops as the elevation increases.

Location	Elevation (feet above sea level)	Boiling point of water (°C)
San Francisco, CA	Sea level	100.0
Salt Lake City, UT	4,390	95.6
Denver, CO	5,280	95.0
La Paz, Bolivia	12,795	91.4
Mount Everest	20,028	76.5



Boiling and Boiling Points

- Increasing the pressure can be used to increase the boiling temperature of water.

- This is how pressure cookers work

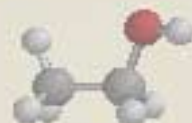
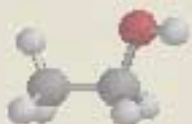
Food cooks quicker in a pressure cooker because water boils at a higher temperature.

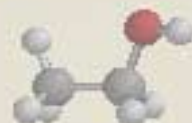
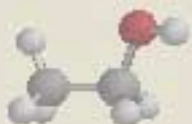
Pressure Above Atmospheric		Boiling Point of Water
<i>Psi</i>	<i>Torr</i>	°C
5	259	108
10	517	116
15	776	121



Sublimation and Melting

- Sublimation
 - Solids also have a vapor pressure.
 - Molecules of the solid can escape to become gas molecules.
 - This process is called **sublimation**.





Sublimation and Melting

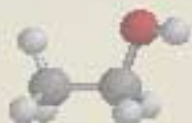
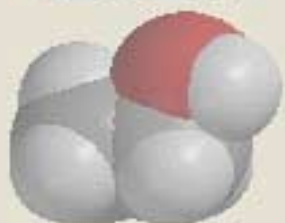
- Melting point
 - Solids also convert to liquids.
 - This occurs when some, but not all of the interactions holding the molecules together in the solid state are disrupted by kinetic energy.

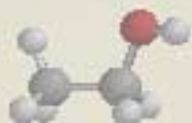
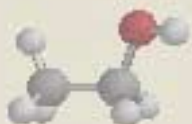
The remaining intermolecular interactions still hold the molecules next to each other, however, the molecules are able slide past one another. interactions.



Sublimation and Melting

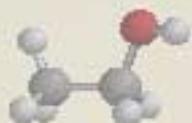
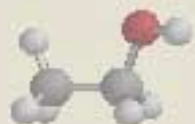
- Decomposition
 - Instead of melting, some solids will decompose when heated.
 - This occurs when the increasing kinetic energy breaks intramolecular (covalent) bonds, instead of the non-covalent intermolecular interactions.





Energy and the States of Matter

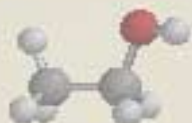
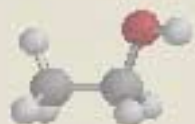
- Energy is added to a substance by heating it
- When heat energy is absorbed by matter it is converted to kinetic energy and the temperature rises
 - Specific heat is the measure of how much heat is needed to raise the temperature of 1 *g* of a substance by 1°C.



Energy and the States of Matter

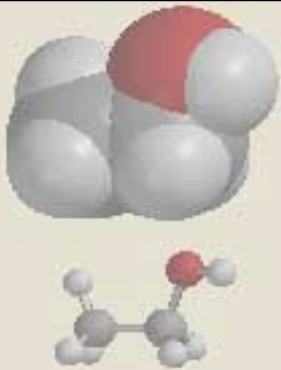
- Specific heats for selected substances:

Substance and state	Specific heat	
	cal/g degree	J/g degree
Aluminum (solid)	0.24	1.0
Copper (solid)	0.093	0.39
Ethylene glycol (liquid)	0.57	2.4
Helium (gas)	1.25	5.23
Hydrogen (gas)	3.39	14.2
Lead (solid)	0.031	0.13
Mercury (liquid)	0.033	0.14
Nitrogen (gas)	0.25	1.1
Oxygen (gas)	0.22	0.92
Sodium (solid)	0.29	1.2
Sodium (liquid)	0.32	1.3
Water (solid)	0.51	2.1
Water (liquid)	1.00	4.18
Water (gas)	0.48	2.0



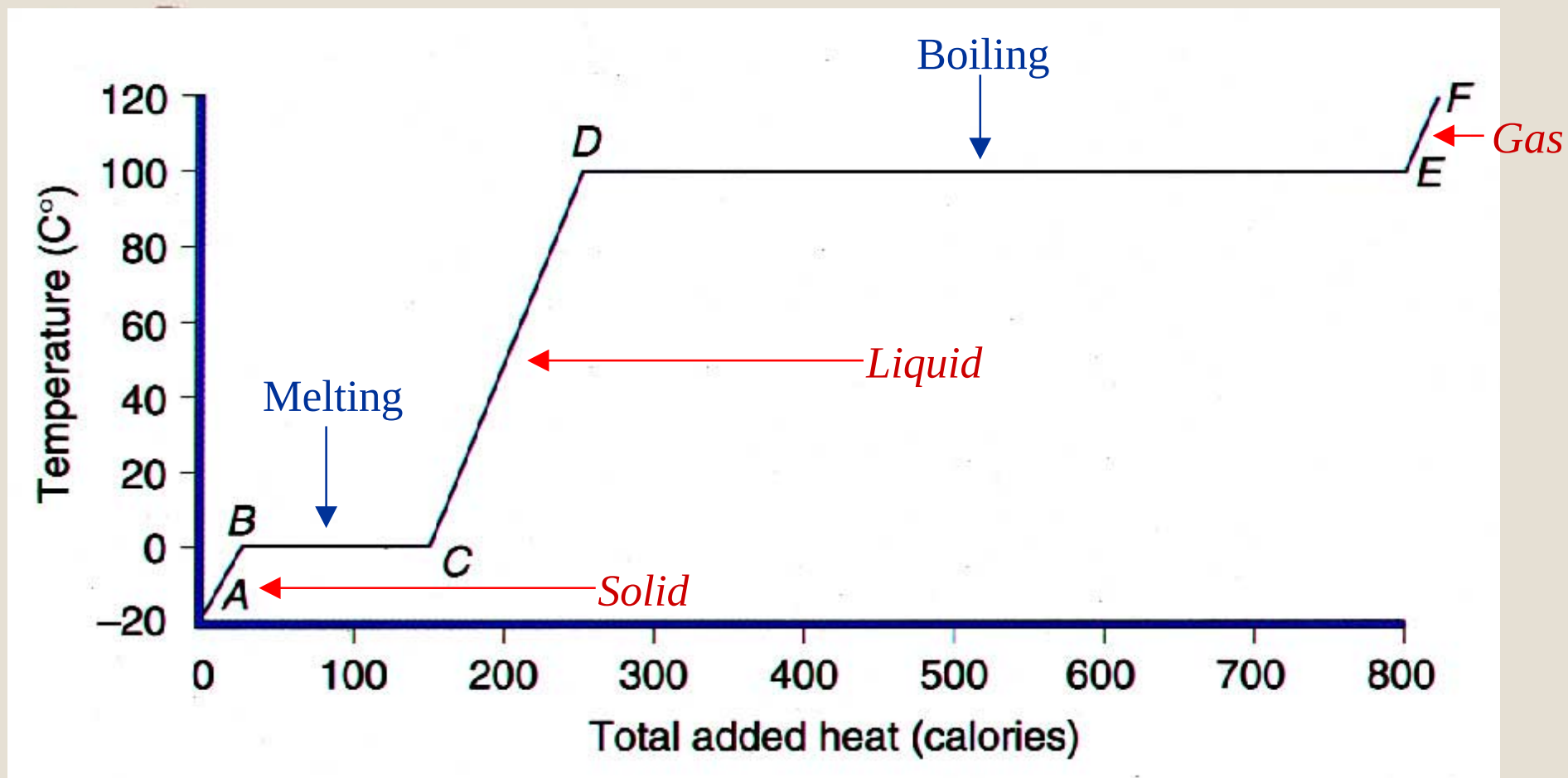
Energy and the States of Matter

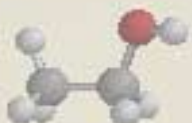
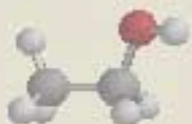
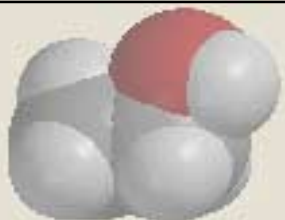
- When heat energy is absorbed and substance changes its state, the energy is converted to potential energy.
 - While the substance is changing states its temperature does not change.



Energy and the States of Matter

- Temperature behavior of matter as it is heated:





Kinetic-Molecular Theory of Matter

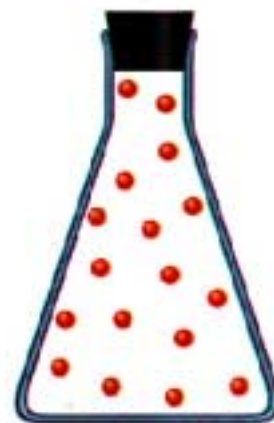
- Disruptive forces are directly dependent on temperature
- Cohesive interactions are independent of temperature.
 - As the temperature increases the disruptive forces become greater and more dominant.



(a) Solid state: The particles are close together and held in fixed positions; they do not need a container.



(b) Liquid state: The particles are close together but not held in fixed positions; they take shape of the container.



(c) Gaseous state: The particles are far apart and completely fill the container.