Ms. Randall
Regents Chemistry
Unit 7: Physical Behavior of Matter Notes

Unit Objectives:

- Distinguish between the three phases of matter by identifying their different properties and representing them with particle diagrams
- Perform simple conversions between Celsius and Kelvin temperature scales
- Differentiate between exothermic and endothermic reactions/changes
- Identify phase changes, and understand how to read a heating or cooling curve
- Define heat, and understand how it varies from temperature
- Solve heat equations
- Solve gas law problems using the following laws: Avogadro's Law, Combined Gas Law, Dalton's Law of Partial Pressures
- State and understand the Kinetic Molecular Theory (KMT)
- Understand the relationship between temperature, volume, and pressure among gases using the following gas laws: Charles' Law, Boyle's Law, Gay Lussac's Law

Focus Questions for the Unit:
- How does particle behavior determine the physical state of a sample of matter?
- How do attractions between particles affect the physical properties of solids, liquids, and gases?
- How do particles in the gas state respond to changes in conditions?

YOU SHOULD BE ABLE TO ANSWER THESE _IN DETAIL_ BY THE END OF THE UNIT

Define the following vocabulary:

Absolute Zero	Kinetic Energy
Avogadro's Law	Kinetic Molecular Theory (KMT)
Boiling Point (Normal)	Matter
Compound	Melting
Cooling Curve	Mixture
Deposition	Potential Energy
Energy	Sublimation
Element	Temperature
Evaporation	Vapor Pressure
Heat	
Heat of Fusion	
Heat of Vaporization	
Heating Curve	

Lesson 1: <u>Chapter Diary 10</u>

Objective: To use a reading comprehension passage to summarize unit concepts

Directions: Directions: After reading the Chapter diary 10 answer the questions in your unit workbook.
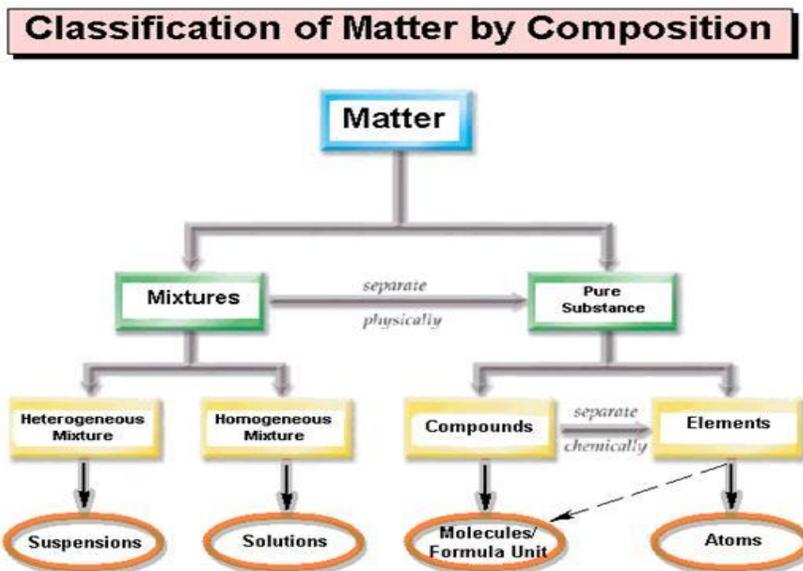
<u>**Lesson 2: Review of Matter**</u>

<u>**Objective**</u>: To recall the vocabulary related to types and phases of matter.

<u>**What have we learned so far:**</u>

Stuff" is made out of particles. How the particles are put together and how they interact with each other and/or energy results in either a physical or chemical change.

Particles are really called atoms and how they are put together causes physical property differences among elements. We organize elements in the Periodic Table based on similarities in their structure.

Elements interact with each other chemically to produce compounds. How these compounds are put together (ionicly or covalently) determines their properties. We can represent these chemical changes with balanced equations and particle diagrams.

Classification of Matter by Composition

Matter

Mixtures — *separate physically* → Pure Substance

Heterogeneous Mixture | Homogeneous Mixture | Compounds — *separate chemically* → Elements

Suspensions | Solutions | Molecules/ Formula Unit | Atoms

Review of Matter

- **Matter**: has definite Mass and occupies a definite Volume.

- **Pure Substance**:
    - Element or Compound

- **Mixture:**
    - Homogeneous
    - Heterogenous

Matter is made of particles which behave in specific ways depending on their structure. Chemist's study these properties to explain observations about the world around them. Now we will focus on the **PHYSICAL PROPERTIES AND CHANGES.**

Recalling Information

1.) STP

Not all substances have the **same phase** at the **same temperature**. Scientists had to develop a term to refer to talk about substances under "**normal**" conditions. It is called **STP**. At STP Oxygen is a gas, while Cu is a solid.
*(SEE REFERENCE TABLE A)*

STP= Temperature ____273K/0 °C_____
Pressure__101.3 kpa/1 atm_____

2.) Phases

Phase	Type of Particle Motion (vibrating, rotating, and/or sliding)	Regular Geometric Arrangement of Particles (position: fixed or not fixed)	Strength of Particle Attractions (weak, strong, or nonexistent)
Solid (s)	vibrating	fixed	strong
Liquid (l)	sliding	not fixed)	weak
Gas (g)	rotating	not fixed)	nonexistent

3.) Phase Changes-Almost all substances can be made to change between the 3 phases, simply by altering the **TEMPERATURE.**
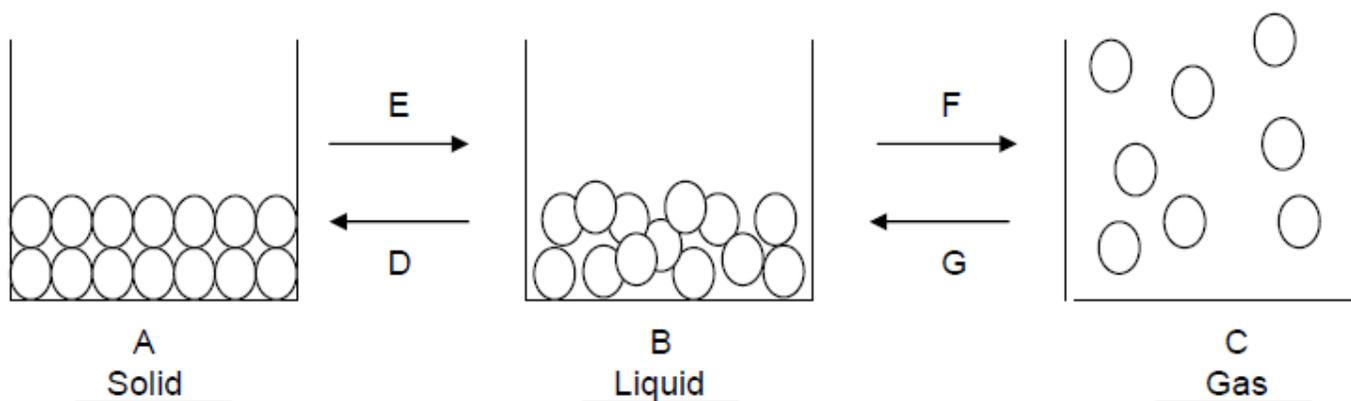
**Properties** can be **physical or chemical**; A **phase change is a physical change**.

Phase Change	Name of Change
Solid to liquid (arrow E)	Melting
Liquid to solid (arrow D)	Freezing
Liquid to gas (occurs only when there is an open container)	Evaporation
Liquid to gas (occurs throughout the liquid) (arrow F)	Vaporization
Gas to liquid (arrow G)	Condensing
Solid to gas	Sublimation
Gas to solid	Deposition

As temperature increases, particle motion increases/and the strength of particle attractions decreases.

Lesson 3: Kinetic Molecular Theory (KMT) Basics

Objective: To relate average kinetic energy and IMF's

I think you would agree that it is an established scientific fact that **particles (molecules, atoms, ions) are in a constant state of random motion.** This was first observed by Robert Brown in 1827, who noticed that pollen grains suspended on the surface of water seemed to "dance about in a random and unpredictable fashion." However, in 1827, molecules and atoms had not yet been discovered. In 1903, Albert Einstein published his first paper, which was an attempt to describe this "Brownian motion" as the response of the pollen grains to collisions with water molecules (molecules and atoms had been discovered in the late 1800's). This was Einstein's introduction to the scientific community. Before this time, he was filling a job as an obscure clerk in the Swedish patent office. This paper later earned him the first of two Nobel Peace Prizes in Physics (the other was for his more famous work in the area of relativity).
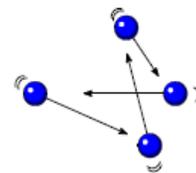
The "KMT" (Kinetic Molecular Theory) attributes the differences in the properties of the 3 phases of matter to differences in particle (molecule or atom) motion. **ENERGY** associated with **MOTION** is referred to as "**KINETIC ENERGY**," so the "**Kinetic Molecular Theory**" can simply be thought of as the "**MOVING MOLECULE THEORY**." *Temperature* is defined as the *Average Kinetic Molecular Energy* of a substance (how fast its particles are moving)

Check this out: **Kinetic Molecular Theory animation**
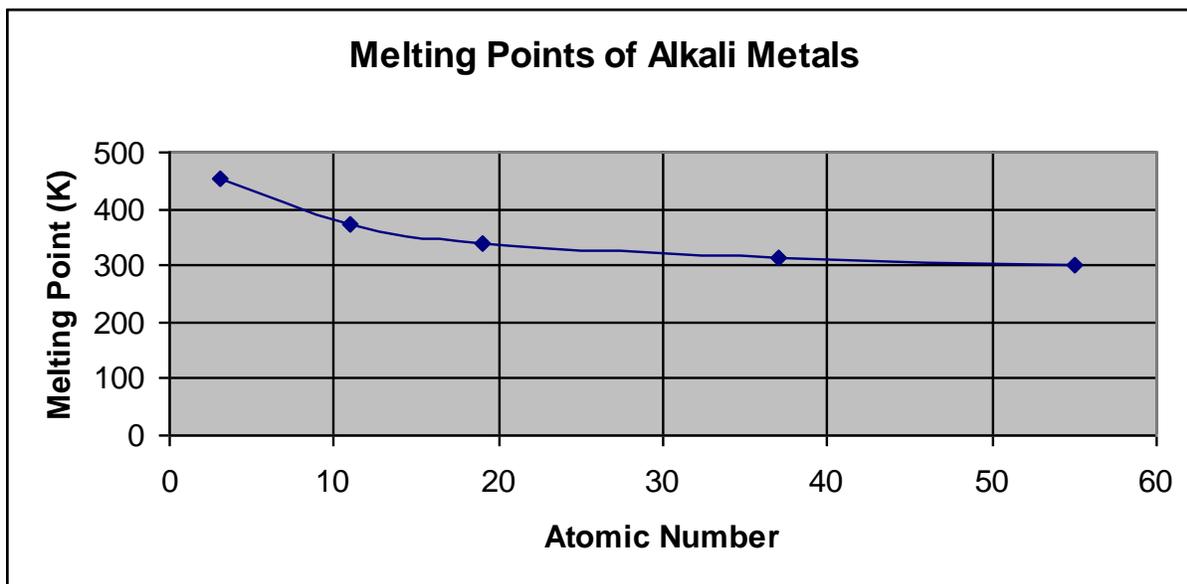
Particle Attractions-IMF!!!

Although it is important to compare the phases of matter from a properties standpoint, we can go a step further and explain why the different phases of matter have the properties that they do from a kinetic molecular theory standpoint. **It really has to do with Intermolecular forces of attraction**, in other words, particles clinging to one another. When comparing a solid to a liquid or to a gas, the relative attraction between particles decreases. This should make sense if you think about the particles in terms of energy.

Particles in a solid have little kinetic energy on average, compared to that of a liquid or a gas. Particles in a solid are simply vibrating and thus remain close together and fixed in position. As heat energy is absorbed by a solid, the particles wiggle and vibrate in place more and more, until finally they have too much energy to stay put. The attractions between the particles in the solid state are weakened as the particles gain energy. At this point the solid starts to melt.

As heat continues to be added, eventually all the solid particles are freed from their positions and can all move and flow past one another, at which point we say the solid has completely melted. As the particles in the liquid state continue to be heated they move faster, as noted by the increase in temperature of the liquid (remember, temperature is a measure of the average kinetic energy of the particles in a sample). As the addition of thermal energy causes the molecules to move faster and faster, eventually there comes a point where they can no longer attract each other at all. When this happens, the particles "escape" from each other, vaporizing into the gas state.

Relationship between IMF and Melting Points

You know from experience that different substances have different melting and boiling points. For instance, Group 1 Elements have different melting points yet they are all in the same family (Alkali Metals). A graph of melting points of the Group 1 elements is below.

Melting Points of Alkali Metals

Melting Point (K) vs Atomic Number

Heating a substance in the solid phase generally causes the temperature to increase. This causes the particles to vibrate and "wiggle" more. As the temperature increases, the particles have a harder time maintaining their attractions for each other. When the forces of attraction are weak enough, melting starts to occur, and now the particles can start rotating and sliding past each other.

The greater the strength of the intermolecular attractions is between the particles of a substance, the higher the melting point temperature is. **The same is true for the boiling point!** It is a direct relationship**!**

At normal atmospheric pressures, some substances can change directly from a solid to a gas. This change is called **sublimation.** This is what occurs when ice cubes shrink in the freezer, or when a wash dries on a clothesline on a cold winter day. Some substances, such as dry ice [$CO_2(s)$], never pass through a liquid phase at standard pressure [101.3 kPa]. A gas can also change directly to a solid. This is called **deposition.**

Lesson 4: Heating and Cooling Curves

Objective: To apply changes in heat and temperature to phase changes

What is heat? What is temperature? How are they different?

**Energy: capacity to do work** (ability of matter to do work)

Kinetic Energy (KE): **energy of motion**/movement associated with temperature change

Potential (AKA Phase) Energy: **energy of position** (stationary or stored energy) associated with phase change, not temperature change

Temperature: measure of the **average kinetic energy of a substance's particles** (Δ TEMP means Δ KE); **does not depend on sample size**; measured in ºC or K

Heat is a **quantity of energy that cannot be measured directly** (as opposed to temperature). It can only be measured as it's transferred from one (hotter) object to another (cooler) object and **does depend** on **sample size** (the larger the sample, the more heat needed to bring it to desired temp). Heat is measured in **Joules (J), Kilojoules (kJ)** or Calories (we will only use J or kJ in this class)

NOTE: Although there is a **direct relationship** between **temperature and heat**, they are not the same thing!
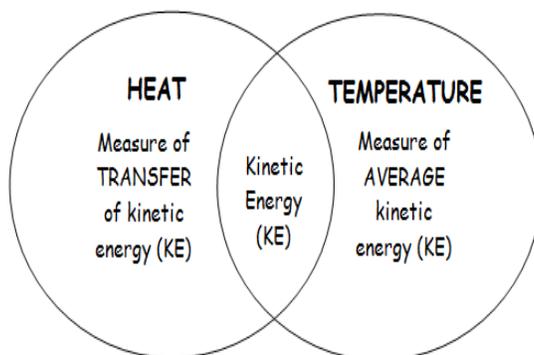

Remember that "heat" is a form of energy, and that it is transferred from Sample A to Sample B (see below) only when the particles in Sample A come into contact with the particles in sample B. The direction of heat flow is determined by the temperature of the two samples.

2. **What is TEMPERATURE?** (Remember that we NEVER use the "f"-word in Chemistry!!! ºF!!!!)

You have probably long thought of temperature as just a measure of how hot or cold an object is. There is another definition that chemists use: **Temperature is equal to AVERAGE KINETIC ENERGY (AVERAGE SPEED)**
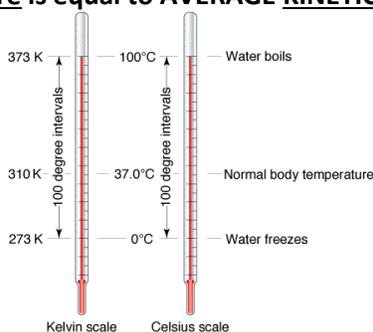
Heating Curves: TEMPERATURE vs. TIME is graphed while a substance is being HEATED at a constant rate.
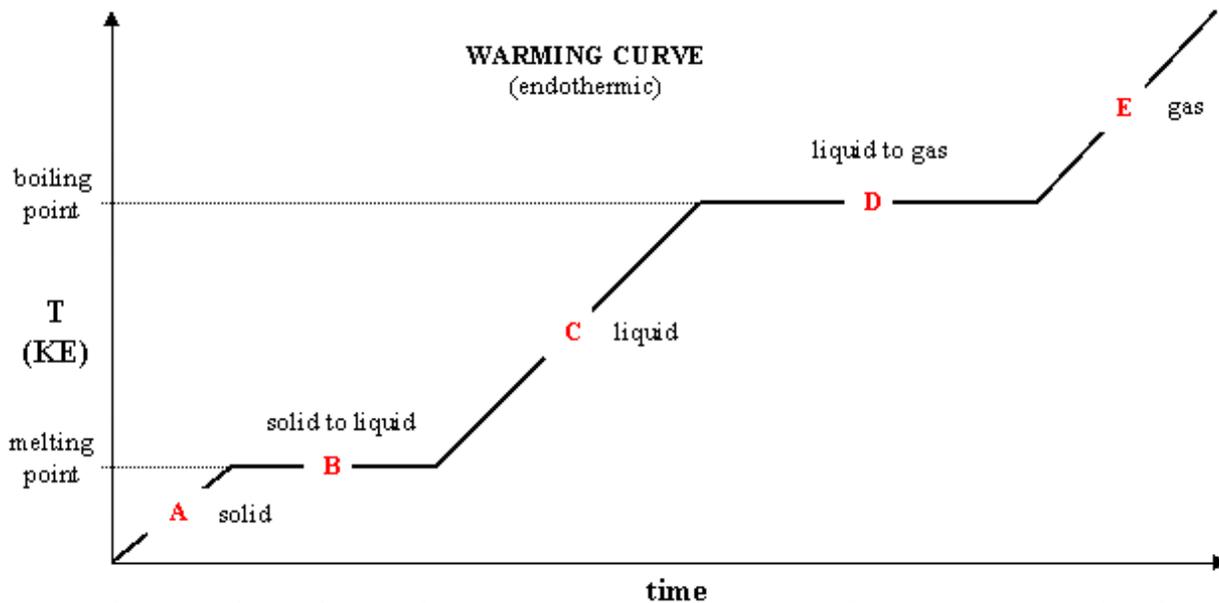
As a substance is heated, its particles begin to move faster and spread apart. The speed of the particles is related to their kinetic energy. The relative position of the particles is related to their potential energy. As solids, liquids, and gases are heated, most of the energy that is absorbed is converted to kinetic energy, and the temperature goes up. But as a substance melts or vaporizes, the particle attractions weaken. As a result, the energy absorbed produces changes in the potential energy of the particles, so the temperature does not change as the phase changes. The **freezing point and the melting point of a substance are the same.**

Cooling Curves: TEMPERATURE vs. TIME is graphed while a substance is being COOLED at a constant rate.
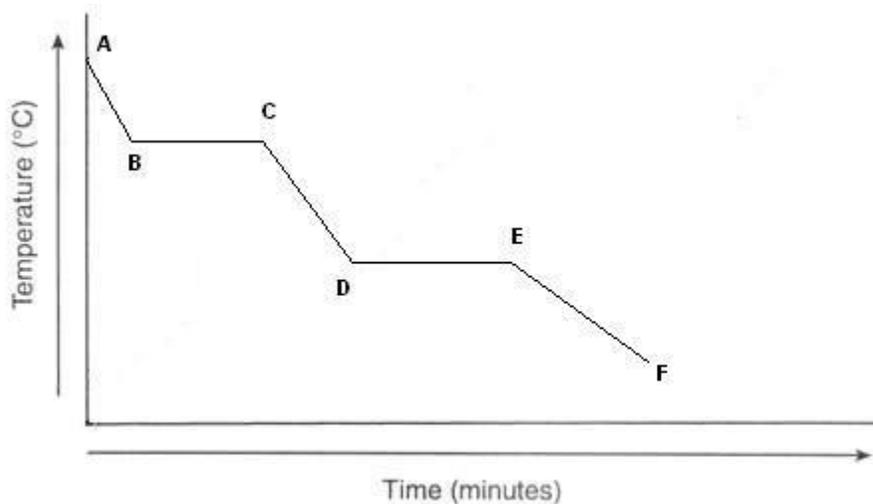

As a substance iscooled, its particles begin to move slower and move closer together. The speed of the particles is related to their kinetic energy. The relative position of the particles is related to their potential energy. As solids, liquids, and gases are cooled, most of the energy is released, and the temperature goes down. But as a substance condenses or freezes, the particle attractions strengthen. As a result, the energy released produces changes in the potential energy of the particles, so the temperature does not change as the phase changes. The **boiling point and the condensing point of a substance are the same.**

MELTING, or "FUSION": We can represent phase changes with particle diagrams. The following particle diagram represents melting. Melting is also known as fusion, so chemists refer to the amount of energy needed to make a substance melt as the "heat of fusion".

melting/fusion

The added energy *decreases* the attractions between particles, causing them to be *less orderly*, but **still attracted** to one another.

BOILING: We can represent phase changes with particle diagrams. The following particle diagram represents boiling. Boiling is a form of vaporization, so chemists refer to the amount of energy needed to make a substance boil as the "heat of vaporization".

vaporization
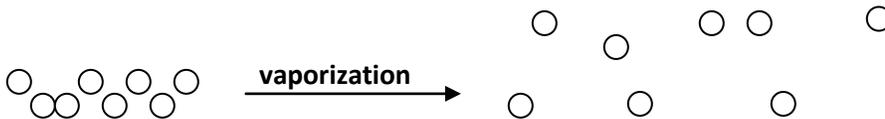
In this case, the added energy *completely overcomes* the attractions between particles, causing them to *become totally dis- orderly* in their conduct, with **no attractions** to one another.
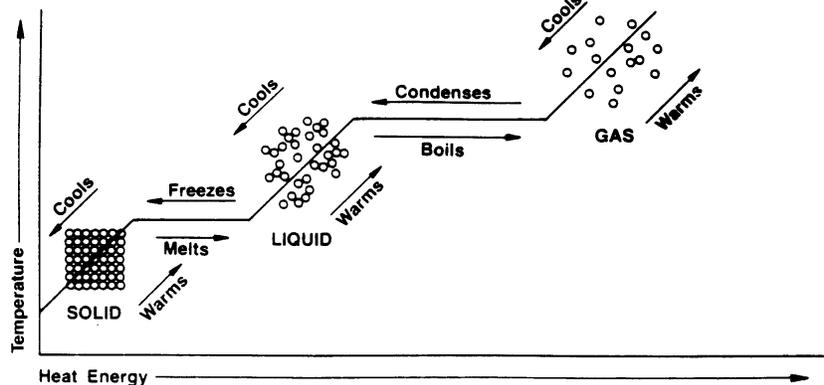
ENERGY

Gas

Liquid

Solid
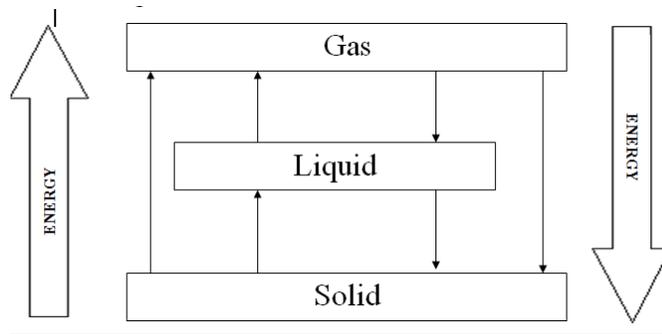
ENERGY

Phase Changes are classified as ENDOTHERMIC or EXOTHERMIC.

ENDOTHERMIC = **system absorbs or takes in heat energy (heating curve)**
EXOTHERMIC = **system gives off heat energy (cooling curve)**

Summary

Temperature

Heat Energy

Cools — Freezes — SOLID — Warms — Melts — LIQUID — Cools — Warms — Condenses — Boils — Cools — GAS — Warms

II. Phase Changes and Heat

A. Endothermic

(s) + heat ⟶ (l)   [melting/fusion]

(l) + heat ⟶ (g) [vaporizing/boiling]

(s) + heat ⟶ (g)   [sublimation]

B. Exothermic

(l) ⟶ (s) + heat

(g) ⟶ (l) + heat

(g) ⟶ (s) + heat

Lesson 5: Measurement of Heat Energy(Potential energy):

Objective: To calculate heat during phase changes and temperature changes

HEAT = Energy transferred due to a difference in temperatures

SPECIFIC HEAT (c) is the amount of heat needed to raise the temperature of 1 gram of a substance by 1°C.

The amount of heat LOST or GAINED in a physical or chemical reaction can be calculated using the following equation (**found in Table T**):

> **q = mcΔT**   q = heat (units = Joules or J)
> m = mass of sample
> c = heat capacity of sample (see Table B for H_2O)
> ΔT = change in temperature (T_f-T_i)

Example 1: How many joules are absorbed when 50.0 g of water are heated from 30.2°C to 58.6°C?
q =?
     m = 50.0g
     c = 4.18 J/g*°C (see Table B)
     ΔT =58.6°C - 30.2°C
     q=(50.0g)(4.18 J/g*°C)(28.4°C)
          q=5935.6J

Why can't we use the equation q = mcΔT to calculate the heat involved in melting or boiling a substance?

Ans: If ΔT = 0ºC, then q = 0 J for a phase change → NOT TRUE

Heat of Fusion (H_f): the amount of heat (or PE) required to change a substance from a solid to a liquid (see Table B)
Hf=334 J/g

Heat of Fusion Equation (see Table T):
q = mH_f

Example: How many joules are required to melt 255 g of ice at 0°C?

q = mH_f
q=(255g)(334 J/g)
q=85170J

Heat of Vaporization (Hv) the amount of heat required to change a substance from a liquid to a gas (see Table B)
 Hv=2260 J/g

Heat of Vaporization Equation (see Table T):
q = mH_v
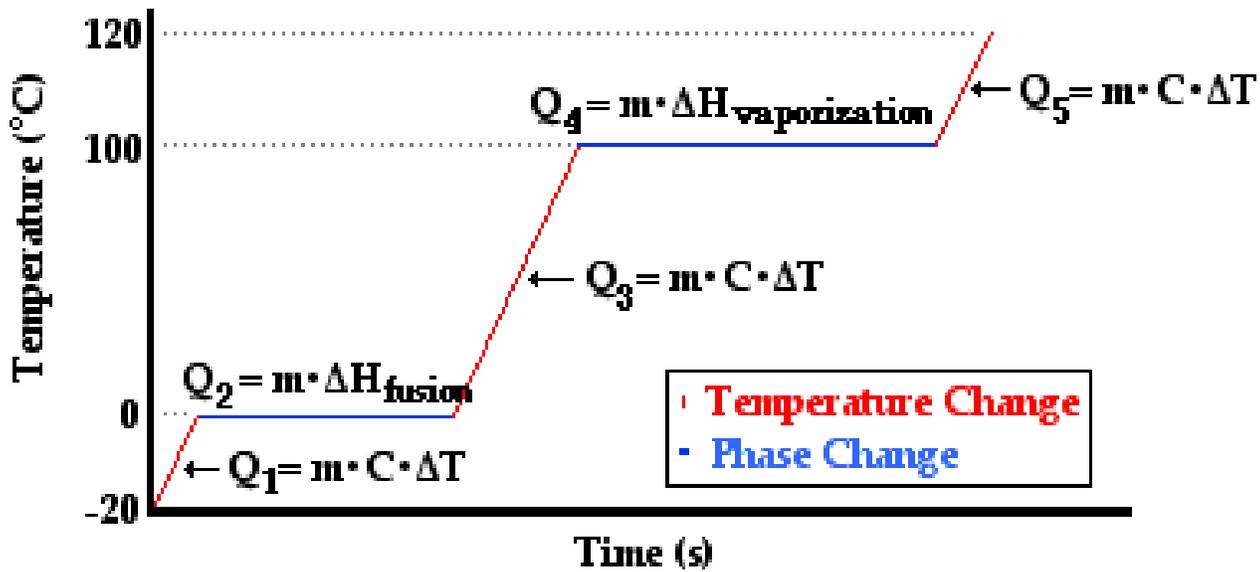Example: How many kilojoules of energy are required to vaporize 423 g of water at 100°C?
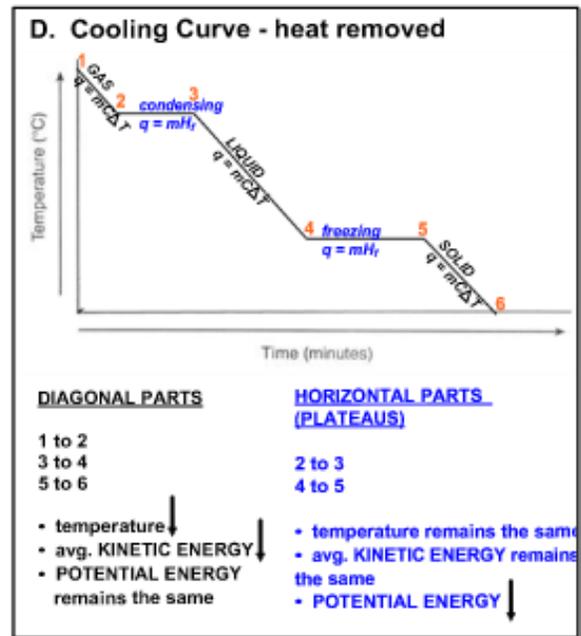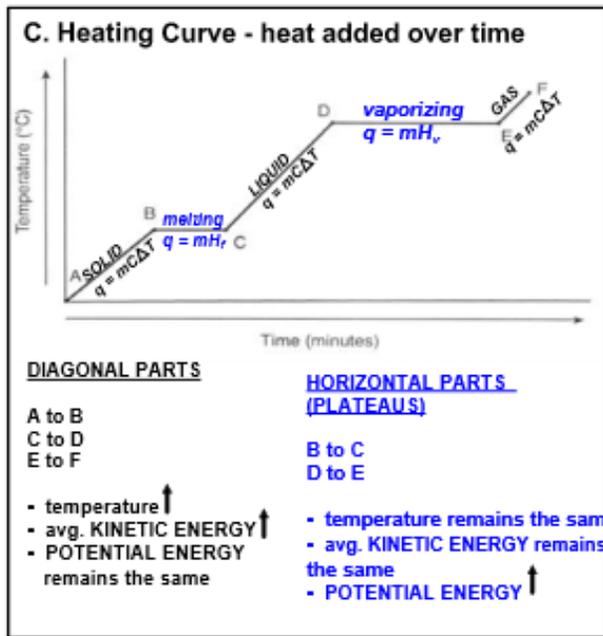**q = mH_v**
q=(423g)(2260J/g)
q=9980J

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How are these formulas applied to a heating curve?

Heating Curve for Water

$Q_4 = m \cdot \Delta H_{vaporization}$

$Q_5 = m \cdot C \cdot \Delta T$

$Q_3 = m \cdot C \cdot \Delta T$

$Q_2 = m \cdot \Delta H_{fusion}$

$Q_1 = m \cdot C \cdot \Delta T$

Temperature Change
Phase Change

Notice that during phase changes there is constant temperature therefore it is not included in the calculation. During heating phases, change in temperature is part of the calculation!

C. Heating Curve - heat added over time

vaporizing
$q = mH_v$

melting
$q = mH_f$

LIQUID
$q = mC\Delta T$

SOLID
$q = mC\Delta T$

$q = mC\Delta T$

DIAGONAL PARTS

A to B
C to D
E to F

- temperature ↑
- avg. KINETIC ENERGY ↑
- POTENTIAL ENERGY remains the same

HORIZONTAL PARTS (PLATEAUS)

B to C
D to E

- temperature remains the same
- avg. KINETIC ENERGY remains the same
- POTENTIAL ENERGY ↑

D. Cooling Curve - heat removed

condensing
$q = mH_f$

freezing
$q = mH_f$

LIQUID
$q = mC\Delta T$

SOLID
$q = mC\Delta T$

$q = mC\Delta T$

DIAGONAL PARTS

1 to 2
3 to 4
5 to 6

- temperature ↓
- avg. KINETIC ENERGY ↓
- POTENTIAL ENERGY remains the same

HORIZONTAL PARTS (PLATEAUS)

2 to 3
4 to 5

- temperature remains the same
- avg. KINETIC ENERGY remains the same
- POTENTIAL ENERGY ↓

Lesson 6: Calorimetry

Objective: To calculate the transfer of heat between substances using the calorimetry formula

Tutorial and lab preparation: Calorimetry pre-lab video

   Energy cannot be CREATED or DESTROYED, only converted from one form to another.
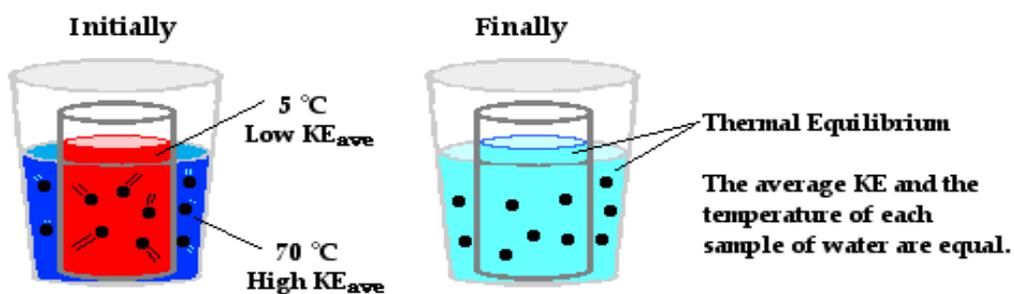   Heat LOST by one substance equals heat GAINED by the other substance.
   A CALORIMETER allows us to measure the amount of heat released by a substance.

   $q_{lost} = q_{gained}$
   Since $q = mC\Delta T$
   $(mC\Delta T)_{lost} = (mC\Delta T)_{gained}$

Lesson 7:   **Vapor Pressure & Boiling Point Temperature**

Objective: To relate vapor pressure to boiling points. To determine the boiling point of a substances using reference table H

Vaporization is different from freezing or melting, because some vaporization occurs at all temperatures of a liquid.

The process of vaporization involves the most active molecules "escaping" from the surface of a liquid.
(It can happen in a solid too – sublimation!)

After vaporization, the molecules that are left behind have a lower average kinetic energy than the original, so the temperature has decreased.

The amount of vaporization is dependent on the temperature of the liquid and on the kind of liquid.
*   The higher the temperature the more vaporization occurs
*   The weaker the intermolecular attractions, the easier molecules escape, the lower the boiling point.
*   The more vaporization at a given temperature, the more cooling you get.

Boiling is when vaporization takes place throughout the bulk of the liquid. The bubbles you see when something is boiling is vapor of the material.

**Vapor pressure** is defined as the "tendency to evaporate
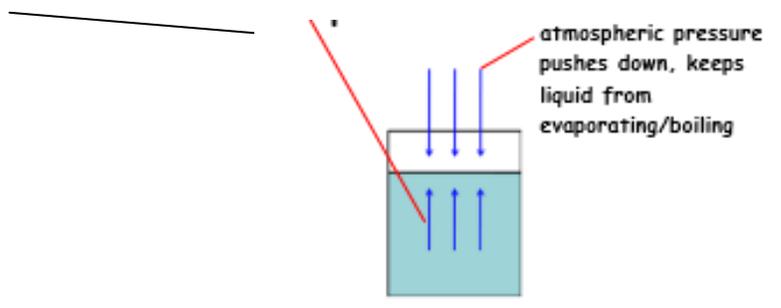
Boiling occurs when the **vapor pressure=atmospheric pressure. The "Normal Boiling point"** is define as the b.p. **at 101.3 kPa or 1 atm.** *Vapor above the liquid causes pressure. The more vapor, the higher the vapor pressure.*
Finally when the vapor pressure above a liquid gets to be the same as the pressure of the atmosphere, that is, the temperature is hot enough, boiling occurs.

Vapor pressure can only be measured if the container is **closed** so that a condition called "***dynamic equilibrium***" can be established between the vapor and the liquid. The relationship between vapor pressure and temperature is **direct**. For any liquid, this means that as the temperature of the liquid increases, the rate of evaporation **increases.** The vapor pressure and atmospheric pressure has to be equal in order for BOILING to occur.

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Table H of your Reference Tables shows the vapor pressures of four common liquids as a function of temperature.
Remember vapor pressure is the UPWARD pressure exerted by a vapor in equilibrium with its liquid. The STRONGER the attractive forces are between liquid molecules, the LOWER the vapor pressure is. Use Reference Table H to determine a liquid's boiling point.

Table H
Vapor Pressure of Four Liquids
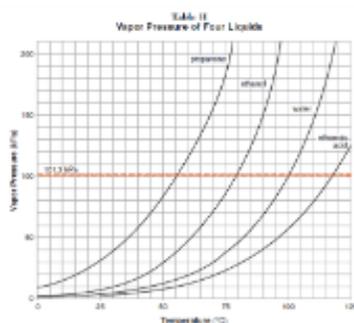
Normal boiling point for

- propanone   57°C ....WEAKEST Intermolecular forces
- ethanol   79°C
- water   100°C
- ethanoic acid 117°C ....STRONGEST Intermolecular forces

- ↑ atm pressure, ↑ b.p.
  (liquid boils at higher temp.)
- ↓ atm pressure, ↓ b.p.
  (liquid boils at a lower temp.)

*At low atmospheric pressure (high elevations)
~ water boils at a LOWER temp.
~ food takes LONGER time to cook

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Vapor pressure and Intermolecular Forces of Attraction (IMF)
b.p., m.p., and evaporation (ALL phase changes) depend on IMF!!

	Strong IMF	Weak IMF
m.p.	high	low
b.p.	high	low
vapor pressure	low "tendency to evaporate"	high "tendency to evaporate"

Lesson 8: Real vs. Ideal gases

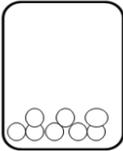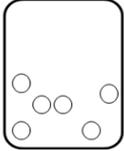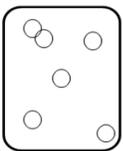**Objective**: To compare the properties of real and ideal gases

    Gases & Entropy   **Entropy** = **disorder, randomness**

            … increases when…   **temperature increases**

                                something dissolves (becomes aqueous)

                                * highest in gases

    Comparing States of Matter: solids, liquids, & gases

• *Particle arrangement: Close, Ordered, vibrate in place* • *Intermolecular forces of attraction: high* • *Shape: definite* • *Volume: definite* • ***Entropy: low***	• *Particle arrangement: Less ordered, flow around eachother* • *Intermolecular forces of attraction: medium* • *Shape: takes shape of container* • *Volume: definite* • ***Entropy: medium***	• *Particle arrangement: Far apart; DISORDER; random movement* • *Intermolecular forces of attraction: low* • *Shape: takes shape of container* • *Volume: explands to fill container* ** COMPRESSIBLE* • ***Entropy: HIGH***

    *Gases are most affected by changes in temperature, pressure, and volume.

  .   **IDEAL vs. REAL Gases**

Ideal Gas	Real Gas
- Imaginary	- Actual gas, what we work with in lab
- Follows the gas laws	- Do not follow gas laws exactly
- Particles are NOT attracted to each other	- Particles DO attract each other (have some intermolecular forces of attraction)
- Particles have NO volume (negligible)	
	- Particles DO have some volume…atomic radii

Gases behave *most ideally* under conditions of HIGH temperature and low pressure
BECAUSE

a. particles are moving faster AND

b. particles are farther apart

Less chance of gas particles attracting each other

Gases *deviate* (stray) from ideal under conditions of low temperature and HIGH pressure
BECAUSE

1. particles are moving slower AND

2. particles are closer together

gas particles will attract each other

H_2 and He are nearly ideal gases because they are the smallest and have the weakest intermolecular forces

Kinetic Molecular Theory (KMT) - model that explains behavior of an ideal gas
Four Assumptions: **(memorize)**

- Gas particles are in continuous, random, straight-line motion
- When gas particles collide, energy is transferred from one particle to another (elastic collisions)
- Gas particles have no attraction to each other (no IMF)
- Individual gas particles have no volume (negligible)

Objective: To select and apply the proper formula to calculate pressure, temperature or volume changes in a gas.

Avogadro's Law: GASES at the same TEMPERATURE, PRESSURE, & VOLUME have the same number of MOLECULES or PARTICLES

CO_2	H_2	Ar
100 torr	100 torr	100 torr
5.0 L	5.0 L	5.0 L
800 K	800 K	800 K

Dalton's Law of Partial Pressure: In a mixture of GASES, the TOTAL PRESSURE of the mixture is the SUM of the PARTIAL PRESSURES of each component gas.

$P_{total} = P_1 + P_2 + P_3 + ...$

Gas Laws

Here's an easy way to remember the relationship between pressure, temperature, and volume. NOTICE how the variables are written from left to right in alphabetical order. Place your finger on whatever variable remains constant, then rotate the variable you want to change up or down. Watch what happens to the third variable as a result.

P   T   V

a. If PRESSURE DECREASES at CONSTANT TEMPERATURE, the VOLUME INCREASES.

b. If TEMPERATURE INCREASES at CONSTANT PRESSURE, the VOLUME INCREASES.

c. If TEMPERATURE INCREASES at CONSTANT VOLUME, the PRESSURE INCREASES.

When either P, T, or V is held constant for a gas:
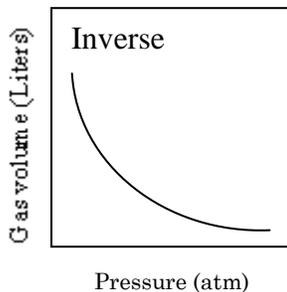
Boyles Law (Constant Temperature):
Example: decreasing P on marshmallow will increase V

$P_1V_1 = P_2V_2$

If the volume of a container is decreased, the pressure inside the container increases. This applies only if the temperature is constant as well as if the amount of gas stays constant. Pressure and volume are inversely proportional. As one goes up, the other goes down.

EXAMPLE: A sample of nitrogen gas occupies 250 mL at STP. What will the new volume of the gas be at 2 atmospheres, with temperature remaining constant?

$V_1 =$ 250mL $V_2 = ?$ P_1 V_1 $=$ P_2 V_2

$P_1 =$ 1 atm $P_2 = 2$ atm $\dfrac{1 \text{ atm } (250mL)}{2 \text{ atm}} = \dfrac{2 \text{ atm } (V_2)}{2 \text{ atm}}$

~~$T_1 =$ 273 K~~   ~~$T_2 = 273$ K~~

(since temperature is constant,

we do not have to use it in the equation)   **125 mL = V_2**

_____

Direct

Gas volume (Liters)

Temperature (K)

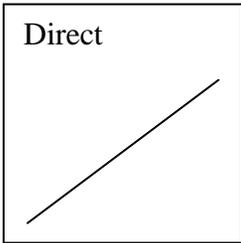### Charles Law (Constant Pressure):
Example: The volume of an ideally behaving gas is 300 L at 227°C.
Example 2: Aerosol cans with gases under high pressure can't be near high temp or contents will expand and the bottle will explode

$\dfrac{V_1}{T_1} = \dfrac{V_2}{T_2}$ **always use K!**

Convert ˚C to K   K = ˚C + 273

As temperature decreases, the kinetic energy of the gas molecules decrease (which means they slow down). Volume and Temperature are directly proportional. As one goes up, the other goes up. As one goes down, the other goes down. This happens any time pressure is kept constant.

Example: While the pressure on 670mL of chloine gas remained steady, the temperature was changed from 24˚C to 0˚C. What change in volume resulted?

$V_1 = 670mL$ $V_2 = ?$

~~$P_1 =$~~   ~~$P_2 =$~~

$T_1 = 24 + 273 = 297K$ $T_2 = 0 + 273 = 273K$

(since pressure is constant, we do not have to use it in the equation)

$\dfrac{V_1}{T_1} = \dfrac{V_2}{T_2}$ $\dfrac{670mL}{297K} = \dfrac{V_2}{273K}$

$\dfrac{(273K)(670mL)}{297K} = V_2$

616 mL = V_2

Pressure (atm)

Direct

Temperature (K)

Gay Lussac's Law (Constant Volume)
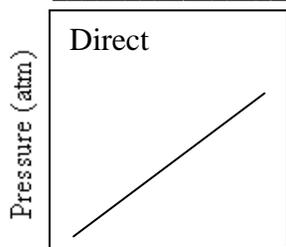
$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

always use K!

(Convert ˚C to K)

K = ˚C + 273

P and T are directly proportional. As one goes up, the other goes up. As one goes down, the other goes down. This happens any time volume is kept constant (i.e. – a rigid container).

Example: A 3.0 mL rigid container is filled with helium gas at a temperature of 150K. The pressure in the container is 167 kPa. If the pressure is increased to 381 kPa, what is the container's new temperture?

~~V_1 = 3.0mL~~   ~~V_2 = 3.0mL~~

P_1 = 167kPa P_2 = 381kPa

T_1 = 150K T_2 = ?

(since volume is constant, we do not have to use it in the equation)

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \qquad \frac{167kPa}{150K} = \frac{381kPa}{T_2}$$

$$\frac{381kPa\ (150K)}{167kPa} = T_2$$

342 K= T_2

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Lesson 10: Combined Gas Laws

Objective: To apply the combined gas law to calculate changes in temperature, pressure or volume of a gas.

Combined Gas Law: (used for GASES ONLY when all <u>THREE VARIABLES</u> for a gas are
 CHANGING – nothing remains constant in this type of problem)

If we combine the three gas laws we get the combined gas law. Look it up on Reference Table T

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Some things to note when using this law…
   1.) The temperature MUST be in Kelvin! If you are given Celsius, convert first!
   2.) "1" = before or initial values; "2" = after or final values.
   3.) "1" units must match with "2" units.

Example: 300 Liters of acetylene gas were measured at a temperature of 273 K and 1 atm. The gas was then
        subjected to a pressure of 3.15 atm and the temperature rose to 281 K. What new volume did the gas
        occupy?

V_1 = 300L V_2 = ? $\frac{P_1\,V_1}{T_1} = \frac{P_2 V_2}{T_2}$

P_1 = 1atm P_2 = 3.15atm

T_1 = 273K T_2 = 281K

(since no variable is constant, all must be used in equation)

$$\frac{(1\ atm\,)(300L)}{273K} = \frac{(3.15atm)(\,V_2)}{281K}$$

$$\frac{(1atm)(300L)(281K)}{(273K)\,(3.15\ atm)} = V_2$$

*cancel out your units → $\frac{84300\ \cancel{atm}\cdot L\cdot\cancel{K}}{859.95\ \cancel{atm\cdot K}}$ = V_2

98.0 L = V_2