Phases of Matter Vocabulary

Word	Definition
Avogadro's Hypothesis	Equal volumes of two ideal gases under the same conditions of temperature and pressure will contain equal number of molecules.
Boiling	The transition of a liquid into a gas at the boiling point
Condensing	The transition of a gas into a liquid at the boiling point.
Deposition	The transition of a gas into a solid.
Equilibrium	The condition that exists when the rates of two opposing changes are equal.
Evaporating	The transition of the surface molecules of a liquid into a gas below the boiling point.
Freezing	The transition of a liquid into a solid at the freezing point.
Gas	A phase of matter characterized by the complete dissociation of matter particles from each other with the distances between the particles very large in comparison to the size of the particles and no attractive forces between them.
Heat of Fusion	The energy required to melt a gram of solid at its melting point.
Heat of Vaporization	The energy required to boil a gram of liquid at its boiling point.
Ideal Gas	A gas in which the molecules are infinitely small and far apart, the molecules travel with a straight-line motion, all collisions have no net loss of energy (elastic), there are no attractive forces between molecules and the speed of the molecules is directly proportional to the Kelvin temperature. Gases are most ideal at high temperature and low pressure.
Liquid	A phase of matter characterized by matter loosely organized yet kept together by intermolecular or ionic attractive forces.
Melting	The transition of a solid into a liquid at the melting point.
Pressure	Force exerted over an area.
Solid	A phase of matter characterized by matter arranged in regular geometric patterns called crystal lattices with only vibration motion, no relative motion.
Specific Heat	The energy required to heat one gram of a substance by one Kelvin.
Sublimation	The transition of a solid into a gas.
Vapor pressure	The pressure exerted by vapor in a vapor-liquid mixture in a closed system at equilibrium.
Vapor-Liquid Equilibrium	A system where the rate of evaporation equals the rate of condensing.

1) Phases and Phase Change (HW: p. 15, 16)

Essential Question: How does the proximity of atoms or molecules to each other affect the properties they exhibit?

Positive and negative charges attract each other. Attractive forces between molecules are called **intermolecular attractive forces**. The strength of these forces determines what phase of matter a substance is in at a given temperature. Substances with weak attractive forces (London dispersion) between their molecules tend to be gases at room temperature, and substances with strong attractive forces (ionic) tend to be solids at room temperature. Phases of matter are simply stages of attraction. Gases are made of molecules with no attractive forces, allowing the molecules to fly freely past each other. Liquids are made of molecules with stronger attractive forces, allowing the molecules to flow past each other, but still stay together. Solids are made of molecules or ions with strong attractive forces, which lock the molecules into a crystal lattice where the particles are free to vibrate, but they cannot move relative to each other.

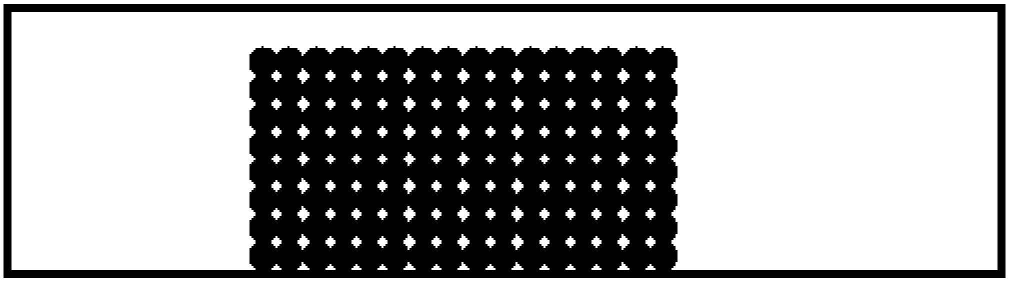
PROPERTIES OF PHASES:

SOLIDS:

a) Molecules, atoms or ions are arranged into a regular, geometric pattern called a crystal lattice.

b) Molecules, atoms or ions vibrate in place. They do not move relative to each other.

c) Definite shape, definite volume.



LIQUIDS:

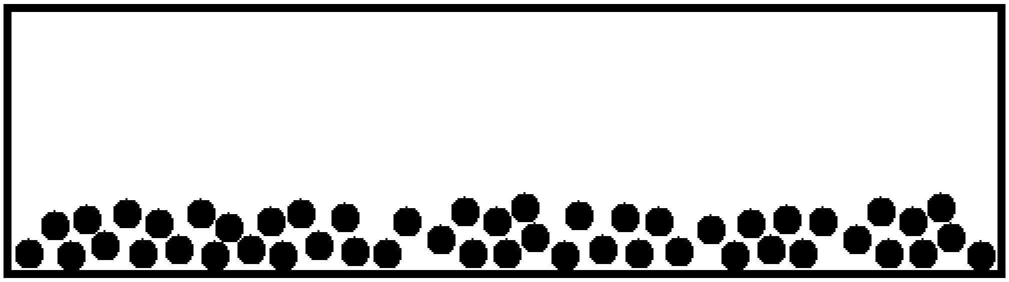
a) Molecules, atoms or ions can flow past each other.

b) Because of intermolecular attractive forces (IMAF), there is resistance to flow, called viscosity.

c) Viscosity increases as temperature decreases and IMAF strength increases.

d) Liquid molecules near the surface can escape into the vapor phase below the boiling point, a process called evaporation.

e) Shape of container (in absence of gravity, they form a perfect sphere), definite volume.



GASES:

a) Gas molecules are extremely far apart compared to the size of the molecules.

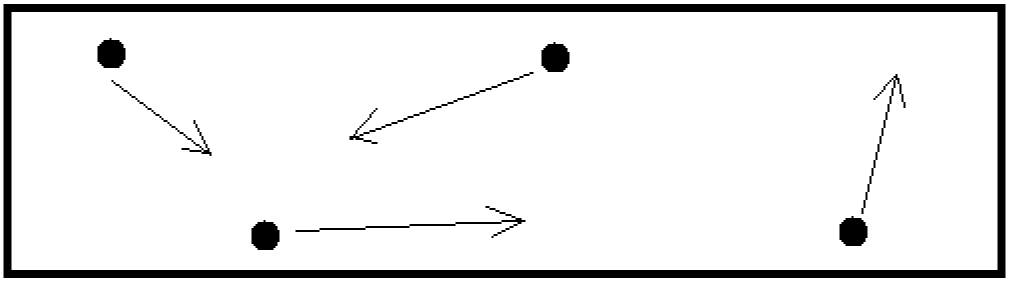
b) Gas molecules travel in a straight line until they collide with something.

c) Collisions are elastic which means they don't lose any kinetic energy in the collision.

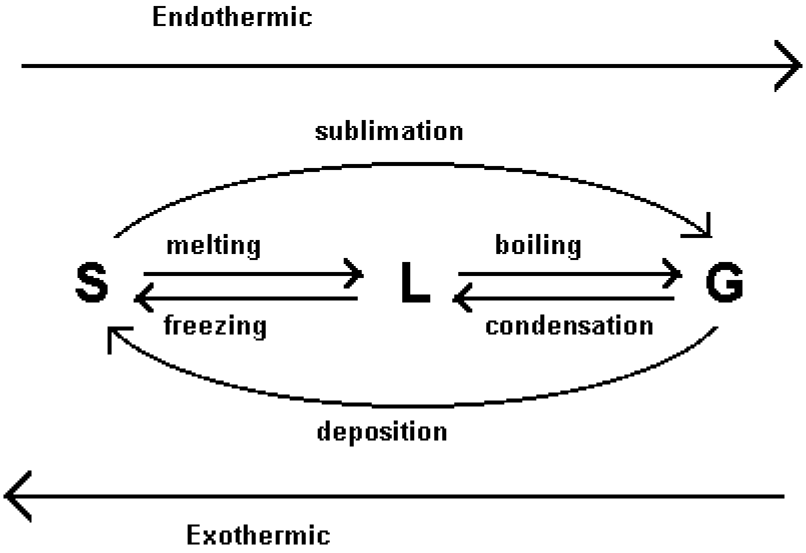
d) Gas molecules move faster when it's hotter (higher Kelvin temperature)

e) The gas phase is the only phase that is affected by changes in pressure.

f) They spread out to take the shape and volume of whatever container they are put into.



Phase Change Diagram:



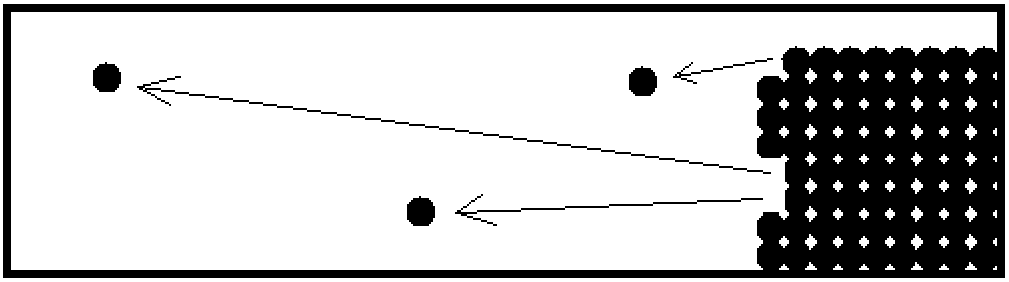
Phase Equilibrium: boiling and condensing both occur at the boiling point (100 ^{o}C for water), freezing and melting both occur at the melting point (0 ^{o}C for water). During the phase change, both phases exist at equilibrium.

EQUILIBRIUM: A condition where the rates of opposing changes are equal. So, a substance at the melting (freezing) point is melting at the same rate that it is freezing.

This is why water melts ABOVE 0^{o}C and freezes BELOW) 0^{o}C. AT 0^{o}C, a mixture of water and ice will not change either way. If you have a sealed flask that contains 20.0 grams of ice and 40.0 grams of liquid water and keep it at 0^{o}C, come back in a day, year, month, century, millennium, eon…there will still be 20.0 grams of ice and 40.0 grams of water. That's equilibrium!

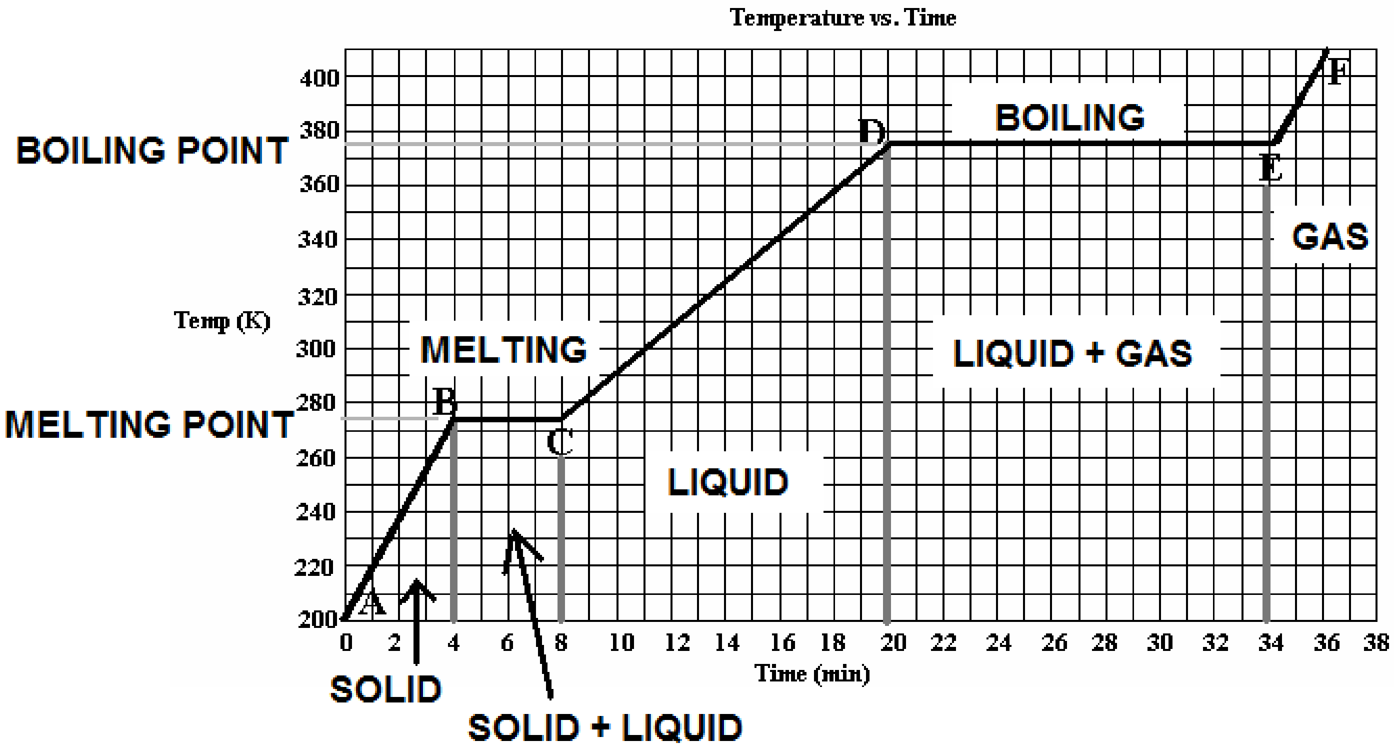
Sublimation: attractive forces between molecules are so weak that heating a solid causes it to go directly into the gas phase. There are two common substances that undergo this change; CO_2 (s), also known as dry ice, and I_2 (s).

Water also undergoes sublimation! Place a fresh batch of ice cubes in the freezer and leave them there for a few months. You will notice that the ice cubes will have shrunk! You also might notice a rime of frost (ice crystals) on the walls of the freezer! Sometimes you have to scrape this frost out, or it builds up thick. What happened is that the surface molecules of water in the ice sublimed (turned into water vapor), hit the very cold walls of the freezer and froze there! When the gas turns directly back into a solid, as happens here (and on your car's windshield on very cold mornings), it is called **deposition**.



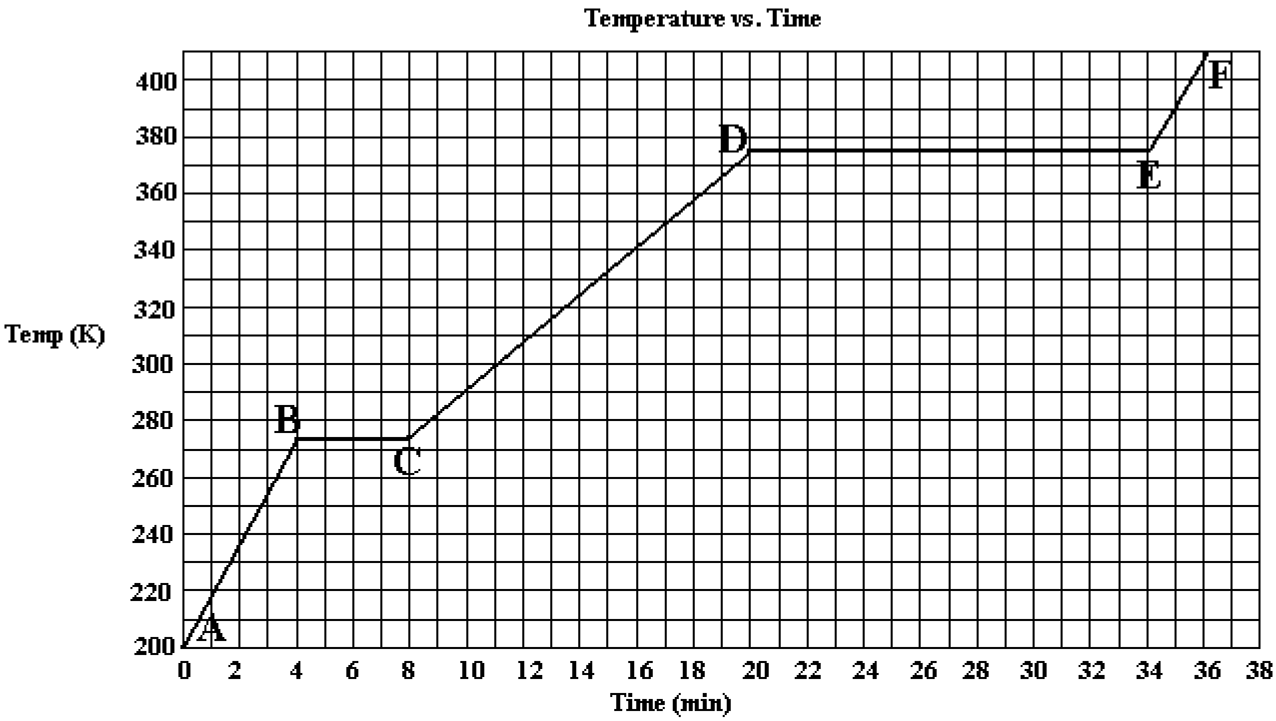
Phase Change Diagrams (Heating Curve):

When a solid (**A**) is heated, its temperature increases until the melting point is reached (**B**)(for H_2O: 0°C or 273 K). The solid then absorbs potential energy as it melts (BC). During the melting process, the temperature remains constant until all of the solid has melted (**C**). Once the substance is completely in the liquid phase, the temperature will increase as the liquid is continuously heated (**CD**). Once the boiling point has been reached (**D**) (for H_2O: 100°C or 373 K), the temperature will remain constant as the substance turns from a liquid into a gas, a process called boiling (or vaporization)(**DE**). Once the substance is entirely in the gas phase (**E**), the temperature will increase as heat is continually applied (**EF**). The phase changes in this scenario are both **endothermic**, because heat is being constantly applied.



Heating Curve For Water

This graph shows what happens to the temperature of a sample of H_2O (s) as it is heated from 200 K to 420 K.



The melting point of water is 0 °C, which is 273 K. The boiling point of water is 100°C, which is 373 K.

How many minutes pass from the first appearance of the liquid phase until the substance is entirely in the gas phase?

The liquid first appears at **B** (4 minutes). It is entirely in the gas phase at **E** (34 minutes). 34-4 = **30 minutes**

How many minutes will it take for this substance to undergo melting?

The solid starts to melt at point **B** (4 minutes), and it is finished melting at point **C** (8 minutes). 8-4 = **4 minutes**

For how many minutes is the water completely in a phase made of a crystal lattice?

The substance is a crystal lattice in the **solid** phase, from **A** (0 minutes) to **B** (4 minutes). 4-0 = **4 minutes**

What line segment represents when H_2O is both in the liquid AND the gas phase?

The substance is in both the liquid and the gas phase during **boiling**, which takes place from **D** (20 minutes) to **E** (34 minutes). 34 – 20 = **14 minutes**

For how many minutes is the water completely in a phase with no definite shape or volume?

The **gas** phase has no definite shape or volume. The substance is completely in the gas phase from **E** (34 minutes) until **F** (36 minutes). 36-34 = **2 minutes**

How many minutes will it take for the water to boil, once the boiling point temperature has been reached?

The substance reaches the boiling point at **D** (20 minutes). Boiling takes until **E** (34 minutes). 34-20 = **14 minutes**

2) Energy Required for Phase Change (HW: p. 17)

Essential Question: How does the defroster in your car work?

1) Melting and boiling are endothermic because energy has to be constantly added in order for the change to be completed.

2) Condensing and freezing are exothermic because energy has to be constantly removed in order for the change to be completed.

HEAT OF FUSION: The amount of heat energy needed to melt one gram of solid at the substance's melting point. For water, the heat of fusion is 334 J/g. You can find this information on Reference Table B. In order to calculate the heat (in joules) needed to melt a sample of a substance, use the equation **$q = mH_f$**. You can find this equation on Reference Table T. If you find that you have frost on your windshield in the morning, turn on the defroster. It will heat the frost to the melting point, and then melt the ice by adding heat of fusion. The same thing is true of your microwave if you are using it to defrost frozen food. The microwave heats the food to the melting point and then adds heat of fusion to melt it so that you can use the food item to make yourself some dinner!

Heat of fusion can also be used for determining how many joules are required to freeze a liquid at the freezing point (which is the same temperature at the melting point). Use the same calculation. The only difference is that when you melt a substance, you have to ADD the heat of fusion to the solid, and when you freeze the substance, you have to REMOVE heat from the liquid. When you put water into a freezer, the freezer removes heat from the water to get it to the freezing point, then removes heat of fusion to turn it all into ice.

How many joules does it take to melt 100. grams of water at its melting point?

$q = mH_f$ = 100. grams X 334 J/gram = **33 400 Joules**

How many joules does it take to freeze 50.0 grams of water at its freezing point?

$q = mH_f$ = 50.0 grams X 334 J/gram = **16 700 Joules**

HEAT OF VAPORIZATION: the amount of heat energy needed to boil one gram of liquid at the substance's boiling point. For water, the heat of vaporization is 2260 J/g. You can find this information on Reference Table B. In order to calculate the heat (in joules) needed to boil a sample of a substance, use the equation **$q = mH_V$**. You can find this equation on Reference Table T. If you have a solution and need to recover the solute dissolved in it, put the solution in an evaporating dish and heat it up over a Bunsen burner or hot plate. Once the boiling point of water has been reached, heat of vaporization boils off all of the water, so you are left with pure solute.

Heat of vaporization can also be used for determining how many joules are released when a gas condenses back into a liquid. Use the same calculation. Which causes greater injury, having boiling water at 100°C splashed on you, or steam at 100°C? The water is hot and your skin absorbs that heat…but the steam is also hot, and as it hits your skin, it releases heat of vaporization as well, causing greater injury. Ask to see my steam burn scar sometime.

Heat of vaporization is absorbed from your skin by sweat as it evaporates, which removes heat from your skin, leaving you feeling cooler. It is our own personal, if slightly gross, air conditioning system.

How many joules are required to boil 100. grams of water at its boiling point?

$q = mH_V$ = 100. grams X 2260 J/gram = **226 000 Joules**

How many joules does it take to condense 50.0 grams of water at its boiling point?

$q = mH_V$ = 50.0 grams X 2260 J/gram = **113 000 Joules**

3) Gases and Pressure (HW: p. 18-20)

Essential Question: What assumptions do we have to make when we set a standard for physical properties?

KINETIC-MOLECULAR THEORY (IDEAL GAS BEHAVIOR)

1) **Gases are made molecules that are extremely tiny and far apart from each other**. This is why gases can be compressed.

2) Gas particles move in a constant straight-line motion.

3) Any collisions a gas particle makes will be elastic that is, there no loss of energy. Each time a gas molecule collides against an obstacle, it bounces off with the same speed that it hit with.

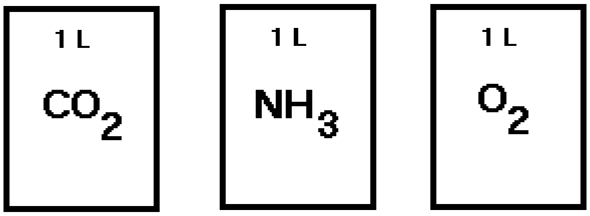
4) There are no intermolecular attractive force between gas particles. All substances have attractive forces, but the distances between the molecules are so great, they don't come into play. Gases with London dispersion forces behave the most like an ideal gas would, and gases with hydrogen bond attractions behave the least like an ideal gas would.

5) The average speed of the particles is directly proportional to the kelvin (absolute) temperature. In other words, the gas molecules move faster when it's hotter.

IDEAL GAS BEHAVIOR: Under conditions of HIGH temperature and LOW pressure, molecules behave more like an "ideal" gas. The smaller the gas molecule is, the more ideal it will behave, therefore, hydrogen and helium are the gases that are the most ideal.

AVOGADRO'S HYPOTHESIS: equal numbers of particles of a gas will occupy equal volume under the same conditions of temperature and pressure. This is because gas molecules will all spread out to an equal degree at the same temperature and pressure. This means that any math we want to do with gases can be applied to all gases equally!

All three gases are at 298 K and 101.3 kPa



They all contain equal numbers of molecules.

DEVIATIONS FROM THE IDEAL GAS LAW - WHERE AND WHY?
This law works well under **standard conditions** of temperature and pressure (STP). (Ref. Table A).

Standard temperature is 0^oC, or 273 K.

Standard pressure is 1.000 atm or 101.3 kPa or 760.0 mmHg (torr).

PRESSURE: force exerted over an area. The gases in the atmosphere exert pressure because of Earth's gravity. The pressure exerted is 14.7 pounds per square inch (psi). This means that over every square inch of surface (including YOU), on average, at sea level, the air molecules in our atmosphere are pushing with 14.7 pounds of force. You are rather lucky in that your internal pressure (and the pressure in your cells) is also about14.7 pounds per square inch. You can feel this pressure if you drive up a mountain or ride in an airplane. As you go up in the atmosphere, there is less air above you exerting pressure. To equalize to this lower pressure, your ears pop!

Pressure is measured in atmospheres (atm), kilopascals (kPa) or millimeters of mercury (mmHg).

CONVERSIONS: 14.7 psi = 1.000 atm = 101.3 kPa = 760.0 mmHg

Convert 2.35 atm to kPa:

(2.35 ~~atm)~~ X (101.3 kPa/~~ atm)~~ = **238 kPa**

Convert 123.4 kPa to atm:

(123.4 ~~kPa~~) / (101.3 ~~kPa~~ /atm) = **1.218 atm**

VAPOR PRESSURE: the pressure exerted by a liquid's vapor in a sealed container at vapor-liquid equilibrium at a given temperature. The vapor pressure of a liquid is not dependent on the mass or volume of the liquid. The vapor pressure of a liquid can be found on **Reference Table H**. **The stronger the attractive forces are between liquid molecules, the lower the vapor pressure is**.

Substances that have high vapor pressures evaporate very quickly. Gasoline is a mixture of liquids with very high vapor pressures, you can even see the gas evaporate as you pump it into your gas tank. Alcohol and acetone (nail polish remover) also evaporate very quickly. Water evaporates fairly slowly, because the hydrogen bonding keeps the liquid molecules more attracted to each other than in the other liquids, meaning the vapor molecules can't escape as easily. Substances that evaporate readily are called **volatile**. Most of these liquids are also quite flammable. The vapor burns even faster than the liquid does. This is why you are not supposed to get back into your car when you are pumping gas until you are done…rubbing your feet on the carpet in your car can cause you to build up a static charge, and if you touch your car, you might cause a spark that could cause an explosion.

BOILING POINT: the temperature at which a liquid's vapor pressure equals the pressure exerted on the liquid by outside forces. Use Reference Table H to determine a liquid's boiling point. Boiling point increases as exerted pressure is increased.

NORMAL BOILING POINT: the boiling point of a liquid under a pressure of 1.000 atmosphere. Substances with higher boiling points have stronger attractive forces holding molecules to each other in the liquid phase, making it require more energy to overcome those attractions and permit boiling.

1) The normal boiling point of water at sea level is 100^oC. At this elevation, the atmosphere is exerting 1.000 atm of pressure.

2) If one takes a ride into the mountains, the amount of air above you decreases, and so does the pressure. Since there is less pressure to keep water in the liquid state, its boiling point decreases. This is why there are special cooking directions for higher elevations. You can't hard-boil an egg when the boiling point is too low!

3) Let's say you have some dried beans and you want to cook them up. To do so, soak them in water for a while, and then add heat. It takes 90 minutes to cook pinto beans (the ones used for refried beans) on the stove at normal atmospheric pressure. Put them in a pressure cooker instead, where the lid is sealed shut so the pressure can just rise, rise and rise), and the boiling point of the water will increase so much that now it will only take 5 to 7 minutes to cook those same beans!

4) On the surface of Mars, the average atmospheric pressure is 0.6 to 1 kPa, depending on where you are. At this pressure, the boiling point of water and the melting point of water are the same! If you were to bring an ice cube on to Mars and heat it, it would undergo sublimation and never form a liquid! That's why any liquid water on Mars must be under the ground, where there is higher pressure to support the liquid phase.

How To Use Table H:

A) What is the vapor pressure of _____ at ____ °C?

What is the vapor pressure of ethanol at 40°C? Start at the 40°C point on the X axis, trace a line up to the ethanol curve, then shoot across and read the vapor pressure off of the Y axis. The vapor pressure is 17 kPa.

B) What is the boiling point of ______ at a pressure of ______ kPa?

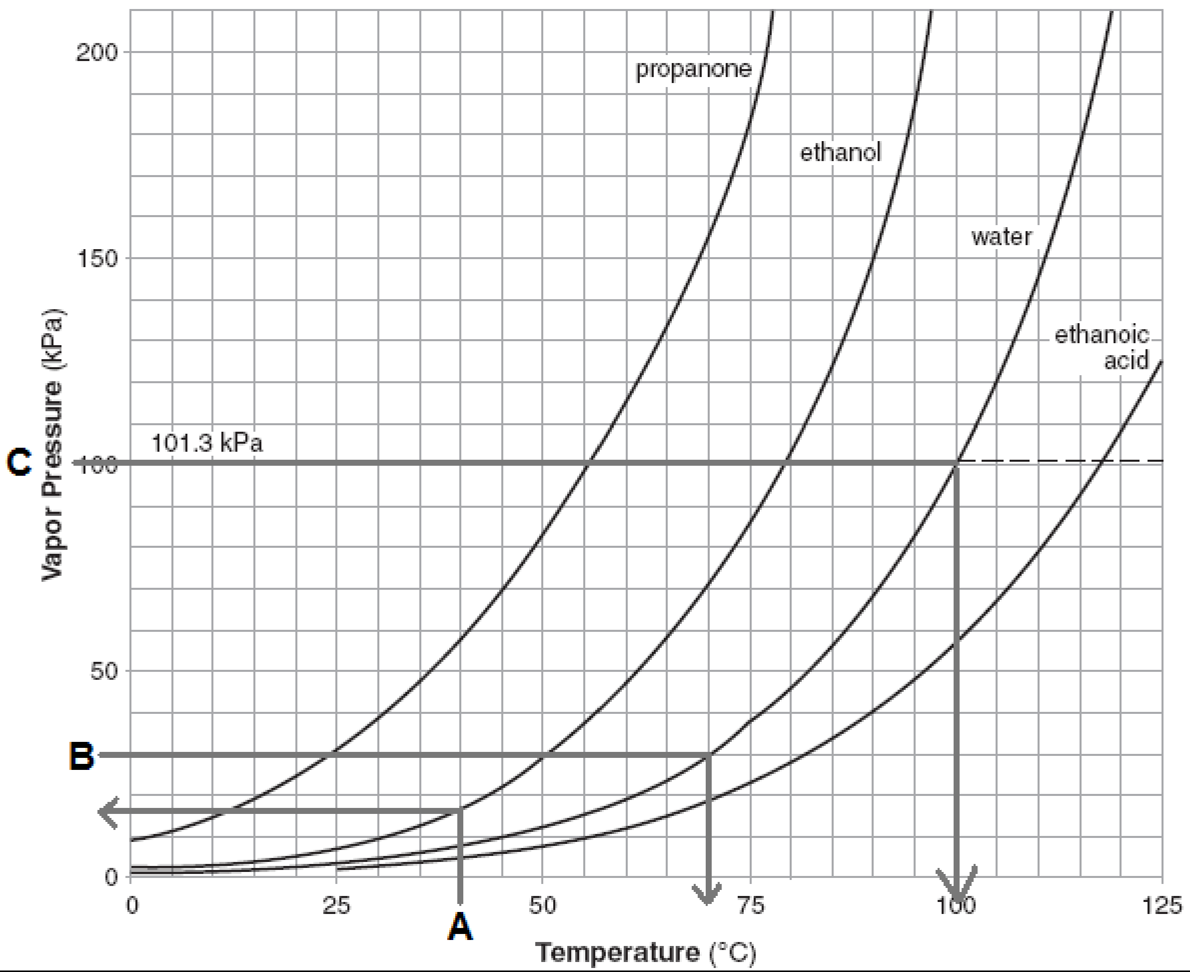
What is the boiling point of water at a pressure of 30 kPa? Start at the 30 kPa point on the Y axis, trace a line to the water curve, then shoot down and read the temperature off the X axis. The boiling point is 70°C.

C) What is the normal boiling point of _____ ?

The normal boiling point is the temperature at which a liquid boils under standard pressure, 101.3 kPa. There is a dashed line going across the table at 101.3 and even labeled "101.3 kPa". Follow the 101.3 kPa line to the curve of your interest (in this case, water), and shoot down to get the boiling point from the X axis. As expected, water has a normal boiling point of 100°C!

Notice that the scales for the X and Y axis are different. Each box on the Y axis is 10 kPa, but every box on the X axis is 5°C.

Table H
Vapor Pressure of Four Liquids



4) The Gas Laws (HW: p. 21-22)

Essential Question: How can models of particle behavior be used to make predictions?

The Gas Laws are relationships between temperature, pressure and volume of a gas. Gas law equations are used to determine what effect changing one of those variables will have on any of the others.

What are the units for pressure, volume and temperature?

Pressure: atm or kPa	Volume: mL or L	Temperature: K

The gas laws are based on one fundamental truth called Avagadro's Hypothesis:
EQUAL VOLUMES OF TWO SAMPLES OF IDEAL GASES CONTAIN EQUAL NUMBERS OF PARTICLES UNDER THE SAME CONDITIONS OF TEMPERATURE AND PRESSURE.

Consider two 4.00 L containers, each at 298 K and 1.00 atm. Container A holds nitrogen gas, Container B holds carbon dioxide gas. If container A holds 2.00 moles of nitrogen gas, how many moles of carbon dioxide must be present in container B?

Since both containers contain equal volumes of gases under the same conditions of temperature and pressure, Avagadro's Hypothesis holds that both containers will contain equal number of molecules, and therefore equal number of moles of molecules. **Therefore, if container A holds 2.00 moles of gas, so must container B**.

This means that all gases behave more or less equally to changes in temperature, pressure and volume, so one equation can be used to describe these changes and applied equally to all gases, if one assumes that they exhibit ideal behavior.

Do equal volumes of gases under the same conditions of temperature and pressure have the same MASS? Why or why not?

No, because each element has it's own unique atomic mass.

To solve gas law problems, follow these easy steps!

1) Get rid of the words! Create a little data table and look through the numbers. If the units are mL or L, it's volume! Pick out V_1 and V_2. If the units are atm or kPa, it's pressure! Pick out P_1 and P_2. If the units are K, it's temperature! Pick out T_1 and T_2. You may not have to pick out all of the variables, as sometimes one of them is constant and you won't have to worry about it.

HINT: If you see the unit ^{o}C, that is temperature, too. In gas laws, temperature MUST be in Kelvin, so simply add 273 and get your T1 and T2 in Kelvin!

2) Write the gas law equation down. Actually write it right there on the page. In the blank spot. Go ahead…

3) Circle the variable you are trying to solve for, and use basic algebra to rearrange the equation. Don't be afraid to ask for help.

4) Substitute the numbers into your newly rearranged equation, and enter them into your calculator. Don't forget to push the "=" or "Enter" button!

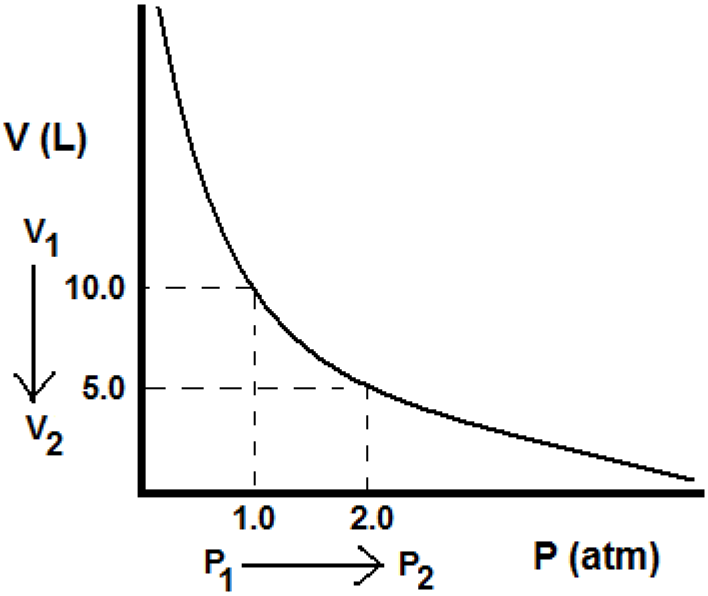
5) Round your answer off. Since this is all multiplication and division, round your answer off to the same number of sig figs as the data you put in that had the fewest number of sig figs! Don't forget to put your unit after your answer!

RELATIONSHIPS BETWEEN VARIABLES OF PRESSURE, VOLUME AND TEMPERATURE

1) Pressure vs. Volume (Constant Temperature): as pressure is increased, volume is decreased (inverse relationship)

BOYLE'S LAW (pressure vs. volume at constant temperature)

$P \times V = K$ $P_1V_1 = K$, $P_2V_2 = K$, therefore $\mathbf{P_1V_1 = P_2V_2}$



A 10.0 L (V_1) sample of gas is trapped in a cylinder with a movable piston at constant temperature. The pressure of the gas in this cylinder is 1.0 atm (P_1). The cylinder is compressed until the pressure doubles to 2.0 atm (P_2). The new volume of the gas will be 5.0 L (V_2). If pressure is doubled, then volume is cut in half.

A sample of gas occupies a volume of 2.00 L at STP. If the pressure is increased to 2.00 atm at constant temperature, what is the new volume of the gas?

Set up a table: (since P_1 is given as STP, look up its value on Table A…it's 1.00 atm!)

$P_1 = 1.00$ atm $V_1 = 2.00$ L $P_2 = 2.00$ atm $V_2 = x$

Then rearrange the Combined Gas Law Equation to solve for the desired variable, omitting the variable that is held constant:

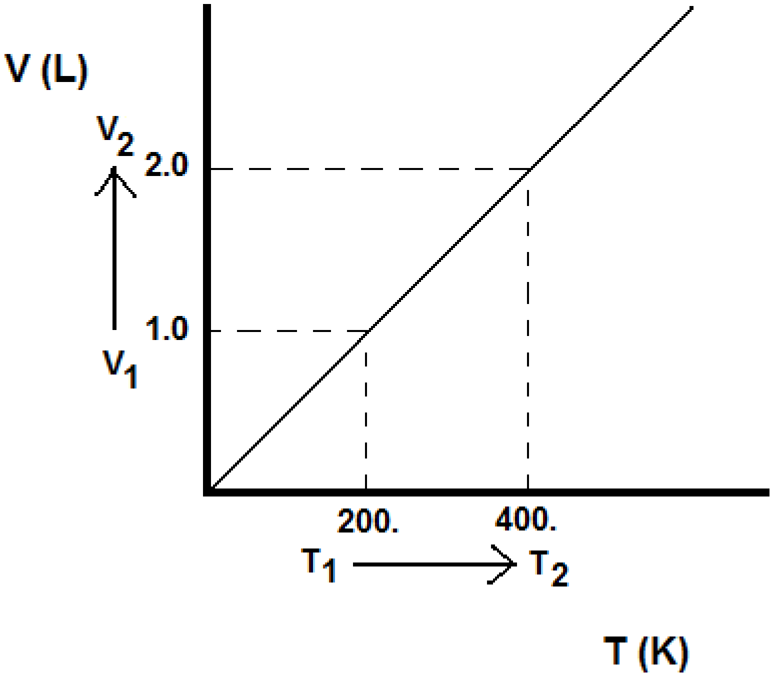
$P_1V_1 = P_2V_2$, solve for V_2 gives a rearranged equation of $P_1V_1/P_2 = V_2$

Then plug and chug! [(1.00 atm)(2.00 L)] / (2.00 atm) = **1.00 L**

2) Volume vs. Temperature (Constant Pressure): as temperature is increased, volume is increased (direct relationship)

CHARLES' LAW (volume vs. temperature at constant pressure)

$V/T = K$ $V_1/T_1 = K$, $V_2/T_2 = K$, therefore $\mathbf{V_1/T_1 = V_2/T_2}$



A 1.0 L (V_1) sample of gas is trapped in a cylinder with a moveable piston at constant pressure. The temperature of the gas in this cylinder is 200. K (T_1). The cylinder is heated until the temperature doubles to 400. K (T_2). The new volume of the gas will be 2.0 L (V_2). If temperature is doubled, then volume is doubled, too.

The temperature MUST be in Kelvin. If it is given in °C, you must add 273 to the temperature before solving the problem.

A sample of gas occupies a volume of 5.00 L at 300. K. If the temperature is doubled under constant pressure, what will the new volume of the gas be?

Set up a table: $V_1 = 5.00$ L $T_1 = 300.$ K $V_2 = x$ $T_2 = 600.$ K(double T_1)

Then rearrange the Combined Gas Law Equation to solve for the desired variable, omitting the variable that is held constant:

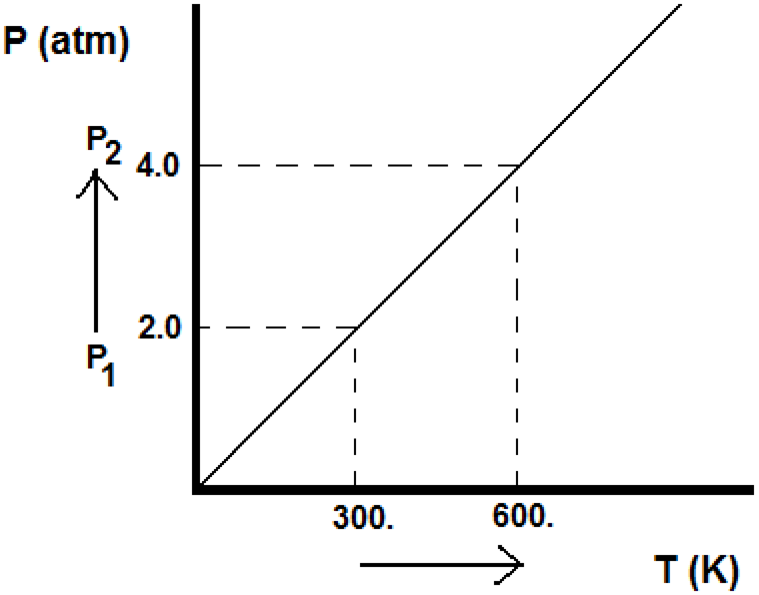
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

, solve for V_2 gives a rearranged equation of $V_1T_2/T_1 = V_2$

Then plug and chug! [(5.00 L)(600. K)]/(300. K) = **10.0 L**

3) Temperature vs. Pressure (Constant Volume): as temperature is increased, pressure is increased (direct relationship)

GAY-LUSSAC'S LAW (pressure vs. temperature at constant volume)
$P/T = K$ $P_1/T_1 = K$, $P_2/T_2 = K$, therefore $\mathbf{P_1/T_1 = P_2/T_2}$



A sample of gas is trapped in a rigid cylinder of fixed volume at 300. K (T_1). The pressure of the gas in this cylinder is 2.0 atm (P_1). The cylinder is heated until the temperature doubles to 600. K (T_2). The new pressure of the gas will be 4.0 atm (P_2). If temperature is doubled, then the pressure is doubled too.

The temperature MUST be in Kelvin. If it is given in °C, you must add 273 to the temperature before solving the problem.

A 10.0 L sample of gas in a rigid container at 1.00 atm and 200. K is heated to 800. K. Assuming that the volume remains constant, what is the new pressure of the gas?

Set up a table: $P_1 = 1.00$ atm $T_1 = 200.$ K $P_2 = x$ $T_2 = 800.$ K

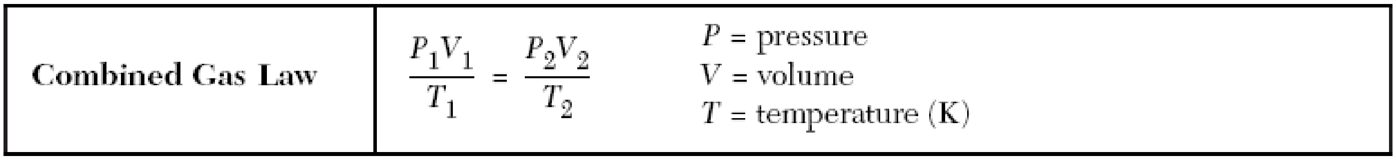
Then rearrange the Combined Gas Law Equation to solve for the desired variable, omitting the variable that is held constant:

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$, solve for P_2 gives a rearranged equation of $P_1T_2/T_1 = P_2$

Then plug and chug! [(1.00 atm)(800. K)]/(200. K) = **4.00 atm**

COMBINED GAS LAW

These three gas laws can be combined to form the **Combined Gas Law**, as found on Reference Table T:



Combined Gas Law	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$	P = pressure V = volume T = temperature (K)

If a variable is constant, it may be ignored. Set up a table for the values of the variables and solve for the unknown. There might not BE anything held constant. In that case, use all of the variables!

A 2.00 L sample of gas at STP is heated to 500. K and compressed to 200. kPa. What is the new volume of the gas?

Set up a table:

P_1 = 101.3 kPa V_1 = 2.00 L T_1 = 273. K P_2 = 200. kPa V_2 = x T_2 = 500. K

Then rearrange the Combined Gas Law Equation to solve for the desired variable:

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$, solve for V_2 gives a rearranged equation of $(P_1V_1T_2) / (P_2T_1) = V_2$

Then plug and chug! [(101.3 kPa)(2.00 L)(500. K)] / [(200. kPa)(273 K) = **1.86 L**

A 2.00 L sample of gas at 1.00 atm and 300. K is heated to 500.K and compressed to a volume of 1.00 L. What is the new pressure of the gas?

Set up a table:

P_1 = 1.00 atm V_1 = 2.00 L T_1 = 300. K P_2 = x V_2 = 1.00 L T_2 = 500. K

Then rearrange the Combined Gas Law Equation to solve for the desired variable:

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$, solve for P_2 gives a rearranged equation of $(P_1V_1T_2) / (V_2T_1)$

Then plug and chug! [(1.00 atm)(2.00 L)(500. K)] / [(1.00 L)(300. K)] = **3.33 atm**

A 2.00 L sample of gas at 300. K and a pressure of 80.0 kPa is placed into a 1.00 L container at a pressure of 240. kPa. What is the new temperature of the gas?

Set up a table:

P_1 = 80.0 kPa V_1 = 2.00 L T_1 = 300. K P_2 = 240. kPa V_2 = 1.00 L T_2 = x

Then rearrange the Combined Gas Law Equation to solve for the desired variable:

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$, solve for T_2 gives a rearranged equation of $(P_2V_2T_1) / (P_1V_1)$

Then plug and chug! [(240. kPa)(1.00 L)(300. K)] / [(80.0 kPa)(2.00 L)] = **450. K**

Student Name:___________________________________ Grades: _____, _____, _____, _____
Phases Heat Gases Gas Laws

1) PHASES AND PHASE CHANGE HOMEWORK

A) Multiple Choice Questions: Place your answer in the space in front of each question.

_____**1) Which sample has molecules that flow past each other, yet are still attracted to each other?**
a) $C_6H_{12}O_6$ (s) b) $C_6H_{12}O_6$ (l) c) $C_6H_{12}O_6$ (g)

_____**2) NaCl (l) is said to be boiling when it undergoes a phase change to**
a) solid b) liquid c) gas d) aqueous

_____**3) What phase of matter has definite volume, but not definite shape?**
a) solid b) liquid c) gas

_____**4) Motor oil viscosity is tested at different temperatures to see how effective it is at providing engine lubrication. As the temperature increases, what happens to the viscosity of the oil?**
a) increases b) decreases c) remains the same

Explain, in terms of *molecular motion* and *average kinetic energy*:

_____**4) Which phase change is exothermic?**
a) LiF (s) → LiF (l) b) LiF (l) → LiF (s) c) LiF (s) → LiF (g) d) LiF (l) → LiF (g)

Explain WHY this change is exothermic:

Explain WHY the other changes are NOT exothermic:

_____5**) What phase of matter has the strongest attractive forces?**
a) solid b) liquid c) gas

Explain why attractive forces between particles of a substance determine what phase a substance is in at a given temperature:

B) A beaker that contains 100. g of substance X in the solid phase is heated constantly over a hot plate at the rate of 1000. J/min, generating the following data:

Time (min)	Temp (K)	Time (min)	Temp (K)	Time (min)	Temp (K)	Time (min)	Temp (K)
0	275	6	291	12	351	18	381
1	279	7	291	13	366	19	381
2	283	8	291	14	381	20	381
3	287	9	306	15	381	21	381
4	291	10	321	16	381	22	390
5	291	11	336	17	381	23	399

1) Draw a graph of the data above, placing the temperature on the Y-Axis, and time on the X-axis**. Label the following information on the graph: where the three phases occur and what and where the phase changes are**.

2) Liquid first appears at minute _____________.

3) The substance is first completely in the liquid phase at minute _____________.

4) The gas phase first appears at minute _____________.

5) The substance is first completely in the gas phase at minute_______________.

6) How many minutes pass from the first moment the liquid phase forms until the substance is completely in the gas phase?

7) For how many minutes is the substance completely in the liquid phase?____________________

8) For how many minutes does the substance exist completely as a crystal lattice?___________________

9) For how many minutes is the substance in a phase that has no definite mass or volume?_________________

10) While the substance is being heated in the liquid phase, what happens to the viscosity of the liquid?___________

11) From minutes 4-8, what is happening to the KINETIC ENERGY of the substance?__________________________

12) From minutes 14-21, what is happening to the POTENTIAL ENERGY of the substance?_____________________

13) Water melts at 273 K and boils at 373 K. Which substance, X or water, has stronger intermolecular attractive forces? Explain briefly.

2) Energy Required for Phase Change Homework

A) Multiple Choice Questions: Place your answer in the space in front of each question.

_____1) Which of the following phase changes requires heat of fusion to accomplish?
a) $H_2O\ (s) \rightarrow H_2O\ (g)$ b) $H_2O\ (g) \rightarrow H_2O\ (l)$ c) $H_2O\ (l) \rightarrow H_2O\ (g)$ d) $H_2O\ (s) \rightarrow H_2O\ (l)$

_____2) Which of the following phase changes is endothermic?
a) $H_2O\ (s) \rightarrow H_2O\ (l)$ b) $H_2O\ (g) \rightarrow H_2O\ (l)$ c) $H_2O\ (l) \rightarrow H_2O\ (s)$ d) $H_2O\ (g) \rightarrow H_2O\ (s)$

3) A mixture of 50.0 g of ice (H_2O (s)) and 30.0 g of water (H_2O (l)) sits in a sealed flask at $0^{o}C$.

_____a) What will happen to the amount of ice in the flask if the mixture is left alone at $0^{o}C$?
a) increase b) decrease c) remain the same

Explain your answer:

_____b) What will happen to the amount of ice in the flask if the temperature of the flask is lowered to $-10^{o}C$?
a) increase b) decrease c) remain the same

Explain your answer:

B) Calculate the number of joules required to (show correct numerical setup):

a) melt 20.0 g of H_2O (s) at $0^{o}C$

b) boil 30.0 g of H_2O (l) at $100^{o}C$

c) freeze 200.0 g of H_2O (l) at $0^{o}C$

d) boil 50.0 g of H_2O (g) at $100^{o}C$

3) Gases and Pressure Homework

A) Multiple Choice Questions: Place your answer in the space in front of each question.

_____1) In which way do real gases deviate from ideal gas behavior?
a) real gas molecules are extremely far apart
b) real gas molecules have attractive forces
c) real gas molecules move faster at higher temperatures
d) real gas molecules travel in a straight line

_____2) Under which conditions will O_2 behave most ideally?
a) 100 K and 1 atm b) 200 K and 1 atm
c) 300 K and 1 atm d) 400 K and 1 atm

_____3) Which gas behaves most like an ideal gas at STP?
a) O_2 b) N_2 c) H_2 d) F_2

What quality of an ideal gas does your choice best match?

_____4) Water boils at 100°C at sea level. At the top of Pike's Peak (14110 feet above sea level), water will boil:
a) at 100°C b) below 100°C c) above 100°C

Explain your choice, in terms of *vapor pressure* and *boiling point*:

_____5) Which substance on Reference Table H has the strongest attractive forces holding its molecules together?
a) propanone b) ethanol c) water d) ethanoic acid

Explain your answer, in terms of *normal boiling point* and *energy needed to break molecules apart from each other*:

_____6) Which sample of gas contains the same number of molecules as a 2.0 L sample of O_2 (g) at 300. K and 100. kPa?
a) 1.0 L sample of H_2 (g) at 300. K and 100. kPa
b) 2.0 L sample of N_2 (g) at 300. K and 150. kPa
c) 2.0 L sample of F_2 (g) at 350. K and 150. kPa
d) 2.0 L sample of Cl_2 (g) at 300. K and 100. kPa

Explain, in terms of *Avogadro's Hypothesis*:

B) Perform the following conversions, show your work in the space provided:

1) 2.0 atm = _________________________________ = ___________kPa

2) 1950. kPa = _________________________________ = ___________atm

C) Why is the atmospheric pressure lower at the top of Mount Everest than it is in the bottom of Death Valley? Explain, in terms of the *quantity of air* above each geographical location.

D) In which location will water boil at a higher temperature… at the top of Mount Everest or at the bottom of Death Valley? Explain, using your answer to (C), above, to support your explanation.

E) Based on Reference Table H, What is the vapor pressure of:

a) ethanol @ 80°C? _______________ kPa

b) propanone @ 20°C? _______________ kPa

c) ethanoic acid @ 110°C? _______________ kPa

F) Based on Reference Table H, what is the boiling point of:

a) ethanol under a pressure of 70. kPa? __________________ °C

b) water under a pressure of 10. kPa? __________________ °C

c) ethanoic acid under a pressure of 120. kPa? __________________ °C

G) Based on Reference Table H, what is the normal boiling point of propanone? __________________ °C

H) Why do cooking directions require you to cook your food longer at higher elevations than at lower elevations?

I) A flask containing nothing other than air is filled partway with water and then sealed. What will happen to the pressure inside the flask as time goes on? Explain, in terms of *vapor pressure*.

J) EXTRA CREDIT!!! You have discovered two Earth-sized planets circling around the star Sigma Cygnii. Preliminary investigations through spectral analysis show that H_2O exists on both planets. Here are some other facts about the two planets:

Planet	**Distance from Sigma Cygnii**	**Average Temperature Range**	**Average Atmospheric Pressure**
Garellia	4.5×10^8 km	-40°C – 20°C	50 kPa
Mongar	7.7×10^7 km	60°C – 105°C	20 kPa

On which of these two planets does water exist in the liquid phase for the greatest temperature range? Explain, using information from Reference Table H to support your hypothesis.

4) The Gas Laws Homework

(all work must be shown for credit)

1) A sample of hydrogen gas has a volume of 1.00 L at a pressure of 100. kPa. If the temperature is kept constant and the pressure is raised to 140. kPa, what is the new volume of the gas?

2) A gas sample occupies 10.0 mL at 1.00 atm of pressure. If the gas is put into a sealed 20.0 mL container and the temperature remains constant, what is the resulting pressure of the gas?

3) When 500. mL of hydrogen gas is heated from 30.0°C to 60.0°C at constant pressure, the volume of the gas at 60.0°C will change to:

4) One sunny day in January, you notice that your front left tire looks a bit low. You inflate it to the recommended 35.0 psi (241 kPa). The temperature that day is a nice 40.0°F (4.4°C). Late that night, you have to run an errand, but you notice the tire doesn't look as full as before. You look at the thermometer, which reads a brutal 12.0°F (-11.1°C). What is the pressure in your tire, assuming volume is constant?

5) A sample of gas at 200.K occupies a volume of 10.0 L at a pressure of 2.00 atm. To what temperature must the gas be raised to double both the volume and the pressure?

6) A 2.00 L sample of oxygen exerts a pressure of 2.00 atm at 273 K. If the temperature is raised to 546 K and the volume decreased to 1.00 L, what will be the final pressure on the gas?

7) A gas occupies 10.0 L of volume at STP. How much volume will it occupy at 2.00 atm and 300. K?

8) A CO_2 fire extinguisher contains 15.5 L of CO_2 (g) under a pressure of 500. atm at 22.0°C. When fired, the pressure of the CO_2 drops to 95.0 kPa, and the volume of CO_2 released is 458 L.

a) What is the temperature of the expelled CO_2?

b) CO_2 sublimes (undergoes sublimation) at a temperature of -78.5°C. What phase does the CO_2 exit as?__________

Review Questions:

_____1) At what temperature will the equilibrium H_2O (l) ⇔ H_2O (g) exist?
a) At the freezing point, 273 K b) At the freezing point, 373 K
c) At the boiling point, 273 K d) At the boiling point, 373 K

_____2) At what temperature will the equilibrium H_2O (s) ⇔ H_2O (l) exist?
a) At the freezing point, 273 K b) At the freezing point, 373 K
c) At the boiling point, 273 K d) At the boiling point, 373 K

_____3) Which phase change is sublimation?
a) CO_2 (l) → CO_2 (g) b) CO_2 (l) → CO_2 (s) c) CO_2 (s) → CO_2 (l) d) CO_2 (s) → CO_2 (g)

_____4) Which substance contains molecules arranged in a regular, geometric pattern called a crystal lattice?
a) $C_6H_{12}O_6$ (s) b) $C_6H_{12}O_6$ (l) c) $C_6H_{12}O_6$ (g) d) $C_6H_{12}O_6$ (aq)

_____5) Which sample has no definite shape or volume?
a) $C_6H_{12}O_6$ (s) b) $C_6H_{12}O_6$ (l) c) $C_6H_{12}O_6$ (g) d) $C_6H_{12}O_6$ (aq)

_____6) What phase of matter has the strongest attractive forces between the molecules?
a) solid b) liquid c) gas

_____7) Turning a solid into a liquid is
a) melting, which is endothermic b) melting, which is exothermic
c) freezing, which is endothermic d) freezing, which is exothermic

_____8) How many joules are required to melt 10.0 grams of water at 0°C?
a) 334 J b) 3340 J c) 2260 J d) 22 600 J

_____9) How many joules are required to boil 10.0 grams of water at 100°C?
a) 334 J b) 3340 J c) 2260 J d) 22 600 J

______10) In what way does a real gas different from an ideal gas?
a) real gas molecules are very small. b) real gas molecules are very far apart.
c) real gas molecules have attractive forces. d) real gas molecules move faster as it gets hotter.

______11) Which of the following samples contain equal numbers of molecules?

Gas	Temp (K)	Pressure (atm)	Volume (L)
CO_2	300	2	4
O_2	250	1	3
CH_4	300	2	4
H_2	250	1	7

a) CO_2 and O_2 b) CO_2 and CH_4 c) O_2 and H_2 d) O_2 and CH_4

_____12) A gas will behave most ideally under what conditions?
a) high T, low P b) high T, high P c) low T, high P d) low T, low P

_____13) Which of the following gases behaves most ideally at STP?
a) O_2 b) CO_2 c) H_2 d) NH_3

Activity: Joules needed to melt a gram of ice

Materials: Styrofoam calorimeter cup, thermometer, 600 mL beaker, 100 mL graduated cylinder

Procedure:

1) Measure out exactly 100.0 mL of hot water in the graduated cylinder.
2) Pour the water into the calorimeter.
3) Measure and record the temperature.
4) Immediately place 2 ice cubes into the water and stir. As the ice melts, replace it.
5) Continue adding ice until the temperature does not drop any more, and no more ice melts. Record the final temperature.
6) Pour the water in the calorimeter into the 600 mL beaker until only the ice remains, using the thermometer as a dam to make sure that no ice gets into the beaker. Discard the ice.
7) Pour 100.0 mL of the water into the graduated cylinder and discard it (this represents the 100.0 mL of water you initially put into the calorimeter).
8) Record the volume of the remaining water (this represents the water that resulted from the melting of the ice) using the same graduated cylinder.

Data:

Initial Temperature of Water in Calorimeter	^{o}C
Final Temperature of Water in Calorimeter	^{o}C
Volume of ice that melted	mL

Calculations: The density of water at 25.0^{o}C is 1.00 g/mL.

1) Determine the mass of the water that you put in the calorimeter.   (density of water = 1.00 g/mL)

2) Determine the mass of ice that melted. (density of water = 1.00 g/mL)

3) Determine the temperature change of the water. (ΔT = |final temperature – initial temperature|)

1) Use calculations 1 and 3 to determine the number of joules transferred from the water in the cup to the ice. ($q = m\ c\ \Delta T$)

5) Determine the heat of fusion of ice in joules per gram of ice.
(Answer to step 4 / answer to step 2)

Conclusions (in complete sentences or show all work, including units and sig figs):

1) The accepted value of the heat of fusion of ice is 334 joules per gram. Determine the percent error of your experiment. Show all work, units and proper rounding.

2) What are some possible sources of error for this experiment? How could they be lessened?

3) Describe the heat transfer between the ice and the water in terms of kinetic and potential energy change.