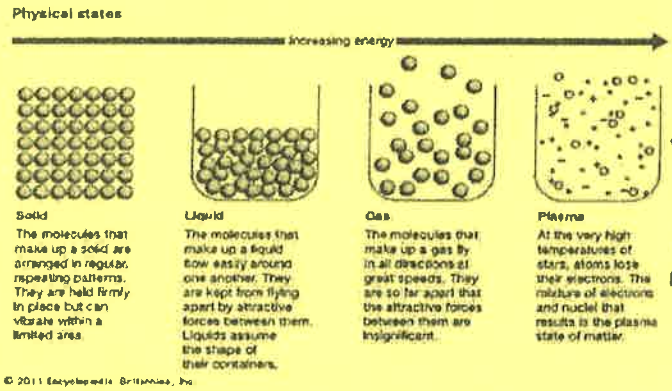
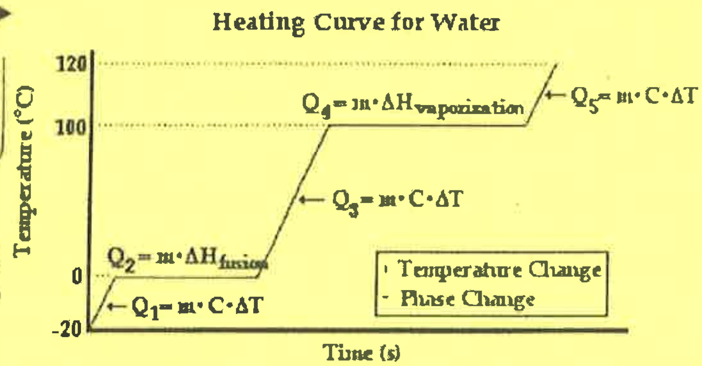
Chapter 10

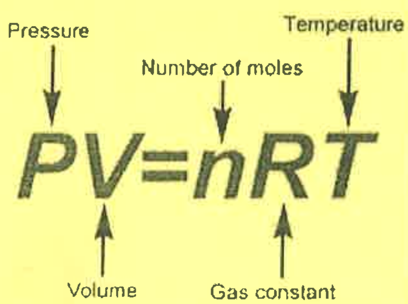
States of Matter





Chapter 11 Gas Laws

Charles's Law:	$V \propto T$ [P and n are held constant] As gas temperature increases, gas volume increases.	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Boyle's Law:	$V \propto 1/P$ [T and n are held constant] As gas pressure increases, gas volume decreases.	$P_1V_1 = P_2V_2$
Avogadro's Law:	$V \propto n$ [P and T are held constant] As the number of moles of gas increase, gas volume increases.	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$
Combined Law:	$V \propto T/P$ [n is held constant] Obtained by combining Boyle's Law and Charles's Law.	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$



Values of the Universal Gas Constant R

Values of R	Units	Values of R	Units
8.314472	$J·K^{-1}·mol^{-1}$	83.14472	$L·mbar·K^{-1}·mol^{-1}$
0.082057	$L·atm·K^{-1}·mol^{-1}$	8.314472×10^{-5}	$m^3·bar·K^{-1}·mol^{-1}$
8.205745×10^{-5}	$m^3·atm·K^{-1}·mol^{-1}$	10.73159	$ft^3·psi·°R^{-1}·lb\text{-}mol^{-1}$
8.314472	$L·kPa·K^{-1}·mol^{-1}$	0.73024	$ft^3·atm·°R^{-1}·lb\text{-}mol^{-1}$
8.314472	$m^3·Pa·K^{-1}·mol^{-1}$	1.98588	$Btu·°R^{-1}·lb\text{-}mol^{-1}$
82.05745	$cm^3·atm·K^{-1}·mol^{-1}$	62.36367	$L·torr·K^{-1}·mol^{-1}$

Name ______________________________ ______________________________ Period ______________

Name____ Date________

Chapter 10 Section 1: The Kinetic-Molecular Theory of Matter

The kinetic molecular theory is based on the idea that :

____.

It can be used to explain the properties of ____________, ______________, and ____________ in terms of the energy of particles and the forces that act between them.

An ideal gas: ____.

1. Gases consist of ____________ numbers of ____________ particles that are ____________ apart relative to their ____________.
2. Elastic collisions ____.
3. Gases are in continuous, ______________, ______________ motion. They possess ________________ ________________, the energy of motion.
4. No force of ____________________ between gas particles.
5. Temperature of a gas depends on the ______________ ______________ ______________ of the particles of the gas. The formula to determine this kinetic energy is

The Nature of Gases:

Expansion: Gases do not have a definite ________________ or ________________.

Fluidity: Gas particles easily glide past one another. Liquids and gases are considered to be ________________.

Density: The density of a gaseous substance at atmospheric pressure is about ______________ the density of the same substance in the liquid or solid state.

Compressibility: Gases can be compressed because

____.

Diffusion and Effusion: Diffusion is the _______________________ mixing of the particles of two substances caused by their ________________ ________________.

Effusion is the process by which gas particles pass through a tiny opening. The rates of effusion of different gases are directly _________________________ to the velocities of their particles. Molecules of low mass effuse faster than molecules of high mass.

Deviations of Real Gases from Ideal Behavior.

A real gas is a gas that does not behave completely according to the assumptions of the kinetic molecular theory. This is because particles of a real gas do occupy space and exert attractive forces on each other. When gases are at high pressures and low temperatures they deviate from ideal behavior. The gases that show the most ideal behavior are the ____________ gases. They are monatomic and nonpolar. Other gases that show ideal behavior are ____________ and ____________. They are diatomic and nonpolar. The more polar a gas the more it will deviate from the ideal behavior. Two common polar gases are ____________, and ______________.

Section 2: Liquids   Properties of liquids and the Kinetic-Molecular Theory

A liquid has a ______________ volume and takes the ____________ of its container.

As in a gas, the particles in a liquid are in constant motion. However, the particles in a liquid are ______________ ________________________ than the particles in a gas are. The ________________ forces are between the particles in a liquid are higher. This is due to the intermolecular forces between the particles. (dipole –dipole, dispersion, hydrogen bonding.) Molecules at the surface of a liquid can overcome the IMF forces and become a gas.

Liquids are more ordered due to the higher IMF forces. The particles are NOT bound in a fixed position according the KE theory. This mobility makes them a fluid. The definition of a fluid is

____.

Most liquids flow downhill due to gravity. However, liquid helium near absolute zero can flow uphill!

Density: Liquids are denser than gases due their close proximity to one another. Most substances are only slightly less dense (%) in a liquid state than compared to a solid state. ____________ is an exception to this rule.

Incompressibility: Liquids have low compressibility compared to gases. At 1000 atm, a liquid may compress 4%. A gas at 1000 atm would compress to 1/1000 of its original volume. Liquids can transmit pressure equally in all directions.

Ability to diffuse: Liquids can diffuse and mix with other liquids. Constant random motion of particles causes this diffusion. Diffusion is ____________ in liquids compared to gases because the particles are closer together. The IMF forces in liquids slows down the diffusion. As the temperature of a liquid increases the diffusion ______________.

Surface tension: This is a force that tends to pull adjacent parts of a liquid's surface together, thereby ________________________ the ________________ area to the smallest possible size. Surface tension is the result of IMF forces between particles. The higher the force ;the higher the surface tension. Why does water have such a high surface tension?

Capillary action: The attraction of the surface of a liquid to the ______________ of a ________________. This explains why water shows a ____________________ in a glass graduate.

Evaporation and Boiling:

The process by which a liquid or solid changes to a gas is called ________________________________. Evaporation is the process by which particles escape the ________________ of a ________________________ liquid.

Evaporation occurs because the particles of a liquid have ________________ kinetic energies. Evaporation removes fresh water from the oceans. This causes the oceans to have a greater concentration of salts. In tropical areas, evaporation occurs faster. Therefore, the oceans are saltier.

Why are you cooled when you sweat?

The physical change of a liquid to a solid by removal of energy as heat is called ________________ or ____________________________.

ime ______________________ Date ______________________ Class ______________________

CONCEPT REVIEW

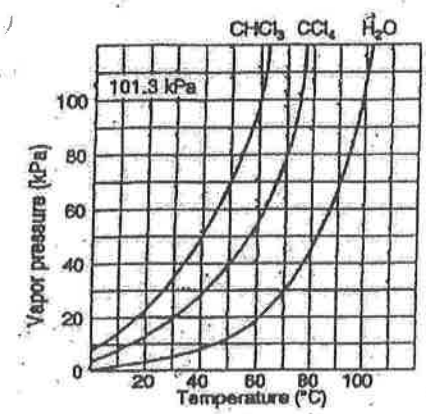


Figure 1

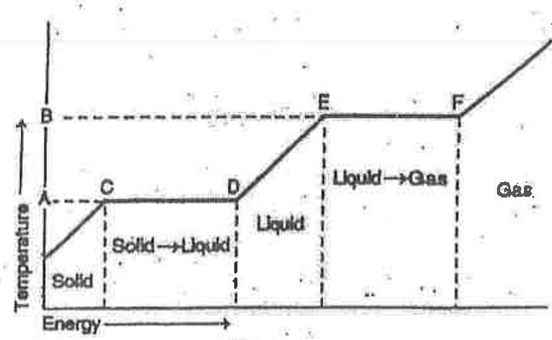


Figure 2

Use Figure 1 to answer the following questions.

1. What will the boiling point of $CHCl_3$ be when the atmospheric pressure exerted on its surface is 101.3 kPa?

____

What is the normal boiling point for CCl_4?

____

3. What is the pressure when water boils at 70°C?

____

4. Which liquid on the graph has molecules that exert the strongest attractive forces on each other?

____

5. Does increasing the temperature of a liquid cause its vapor pressure to increase or decrease?

____

6. Which of the liquids in Figure 1 would be easiest to evaporate?

____

Use Figure 2 to answer the following questions.

7. Identify the point on the graph where each of the following occurs.

____ a. melting begins ____ c. boiling begins

____ b. freezing begins ____ d. condensation begins

Does the substance give off or take in heat as it goes from *D* to *C*?

____

Name ____________________ Date ____________________ Class ____________________

9. Why does the temperature remain constant during a phase change from *E* to *F*?

____

____

10. Explain what happens to volume as water temperature goes from 4°C to 0°C and freezes.

____

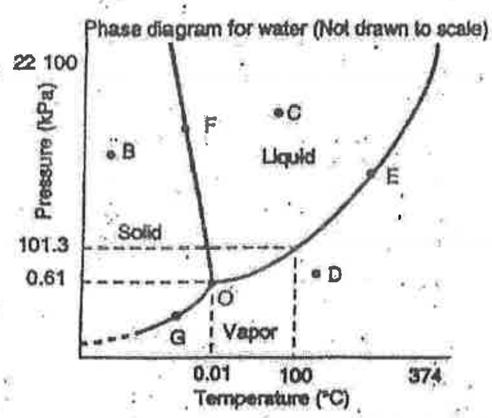


Figure 3

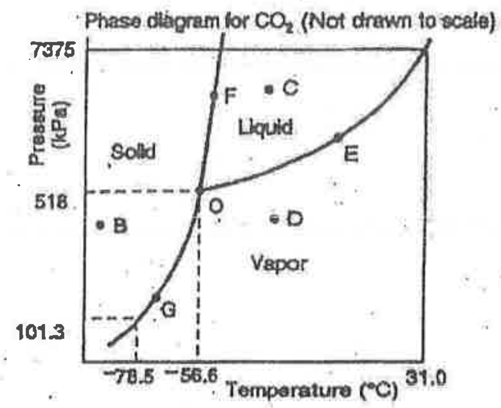


Figure 4

Use the diagrams in Figures 3 and 4 to answer the following questions.

11. Point *O* on the above diagrams represents a temperature and pressure at which all three states exist in equilibrium. What is the point called?

____

12. At which point do only solid and liquid states exist in equilibrium?

____

13. At which point would a liquid be boiling?

____

14. At which point would sublimation occur?

____

15. What is the critical pressure for water?

____

16. What is the critical temperature for CO_2?

____

17. What are the triple point pressure and the triple point temperature for CO_2?

____

18. Refer to the phase diagram for water. What changes in temperature and pressure would be necessary to go from point *D*, vapor, to point *C*, liquid?

____

Name ____ Date______________

Chapter 10 Section 3 Solids

The particles of a solid are more closely packed than those of a liquid or gas. The IMF forces are therefore much more ____________________ in solids. Movement in solids is ___________________________ around fixed points.

Solids have a definite ______________ and ________________. There is very little empty space in between the particles.

Two types of solids: ____________________________ and __________________________.

Most solids are __________crystals____________. A CRYSTAL is a substance in which the particle arrangement is orderly in geometric repeating patterns.

The 3-D arrangement of particles in a crystal is called the _____________________ structure. The arrangement of particles in the crystal can be represented by a system called a ______________________. The smallest portion of a crystal lattice that maintains the 3-D pattern is called a ____________ cell. There are seven types of unit cells. Cubic, tetragonal, hexagonal, trigonal , orthorhombic, monoclinic, triclinic.

Crystals can also be described in terms of the particles in them and their bonding.

- **Ionic:** Cation + Anion Hard, brittle, high M.P, good insulators. ION-ION IMF forces
- **Covalent Network:** Covalent bonding extends throughout a network that includes a very large number of atoms. Diamond, Quartz, Sand are examples. Hard , brittle, VERY high M.P, usually nonconductor or semiconductor. **Metallic crystal:** Metal cations surrounded by a sea of mobile valence electrons. The electrons belong to the crystal as a whole and are free to move throughout the crystal. High electric conductivity
- **Covalent molecular:** Covalently bonded molecules held together by IMF forces. Low melting points, easily vaporized, soft, good insulators. Examples: sucrose, ice crystals

AMORPHOUS solids do not have an orderly geometric repeating pattern. Examples: Glass , plastic, rubber. These materials can be molded in any shape. Amorphous means "without shape".

Melting: Physical change of a solid to a liquid by the addition of heat.

Melting point: temperature at which a solid becomes a liquid. This is the point where the IMF forces between the solid particles are overcome. All crystalline solids have a definite melting point. Amorphous solids DO NOT have a definite melting point. They flow over a wide range of temperatures.

The term ___________________________liquids is used to describe amorphous solids due to the fact that they do not have a definite melting point. This is because they have certain liquid properties even at temperatures where they appear to be a solid. This is because the particles are arranged randomly.

Solids (with the exception of ice) are denser than compared to their liquid state. Solids are considered to be incompressible. Solids also have a very low rate of diffusion (millions of times slower than compared to a liquid.)

Chapter 10 Section 4 Changes of State

________________ is any part of a system that has uniform composition and properties.

________________ is the process by which a gas changes to a liquid.

________________the gas phase of a substance that is a solid or liquid at room temperature.

If the temperature remains constant, and the system is covered. The rate at which molecules evaporate will eventually equal the rate at which the molecules condense. This is called an ______________. Equilibrium is a dynamic condition in which two opposing changes occur at equal rates in a closed system.

Equilibrium vapor pressure is the pressure exerted by a vapor in equilibrium with its corresponding liquid at a set temperature.

How could you increase the vapor pressure of a liquid?

Volatile: A liquid which readily evaporates.

________________ is the conversion off a liquid to the vapor state within the liquid as well as the surface. The boiling point is the temperature at which the equilibrium vapor pressure equals the atmospheric pressure. The lower the atmospheric pressure, the lower the boiling point.

How could you lower the atmospheric pressure and lower the boiling point?

Could you raise the atmospheric pressure and elevate the boiling point?

The temperature at which a substance boils at normal atmospheric pressure (1 atm) is called the ________________ ________________ ______________________________.

To boil a liquid energy must be continually added. The energy is used to break the IMF forces between the molecules. Therefore the temperature of the liquid ____________ ______________ rise during the boiling process.

Molar enthalpy of vaporization, ΔH_v. The amount of energy as heat that is needed to vaporize one mole of liquid at the normal boiling point at constant pressure. The magnitude of the value of ΔHv is proportional to the strength of the IMF forces within the liquid. Water has an exceptionally high ΔH_v.

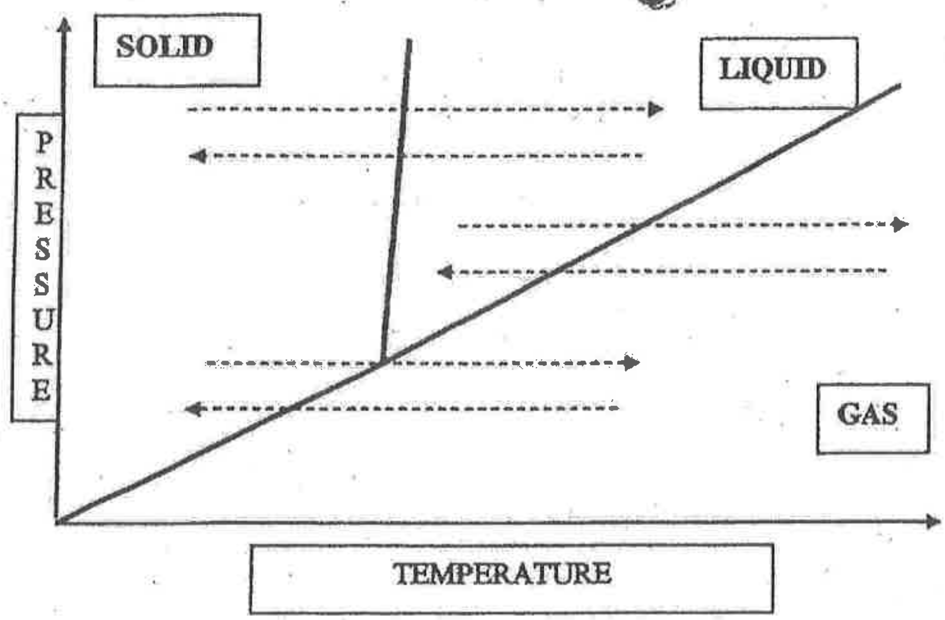
________________is the physical change of a liquid to a solid. Liquid → solid + energy (freezing point = melting point)

The amount of energy as heat required to melt one mole of solid at the solid's M.P. is the molar enthalpy of fusion ΔH_f.

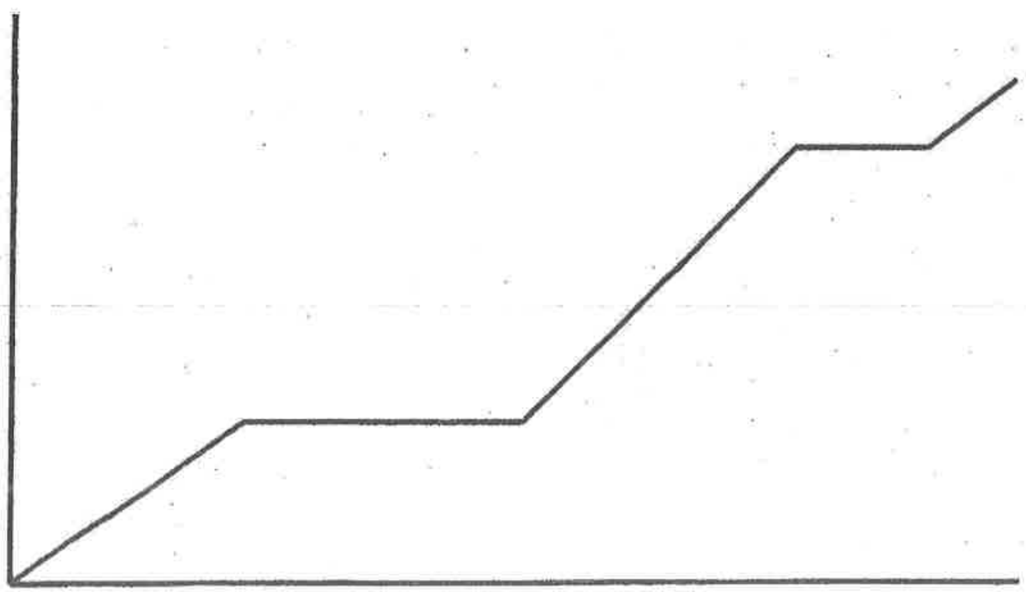
The change of state from a solid directly to a gas is known as ______________. The reverse process is ________________. Dry ice and iodine are two common substances that sublime. Ordinary ice sublimes slowly at temperatures lower than its melting point.

Phase Diagram: a graph of pressure versus temperature that shows the conditions under which the phases of a substance exist. A phase diagram also reveals how the states of a system change with changing temperature or pressure.

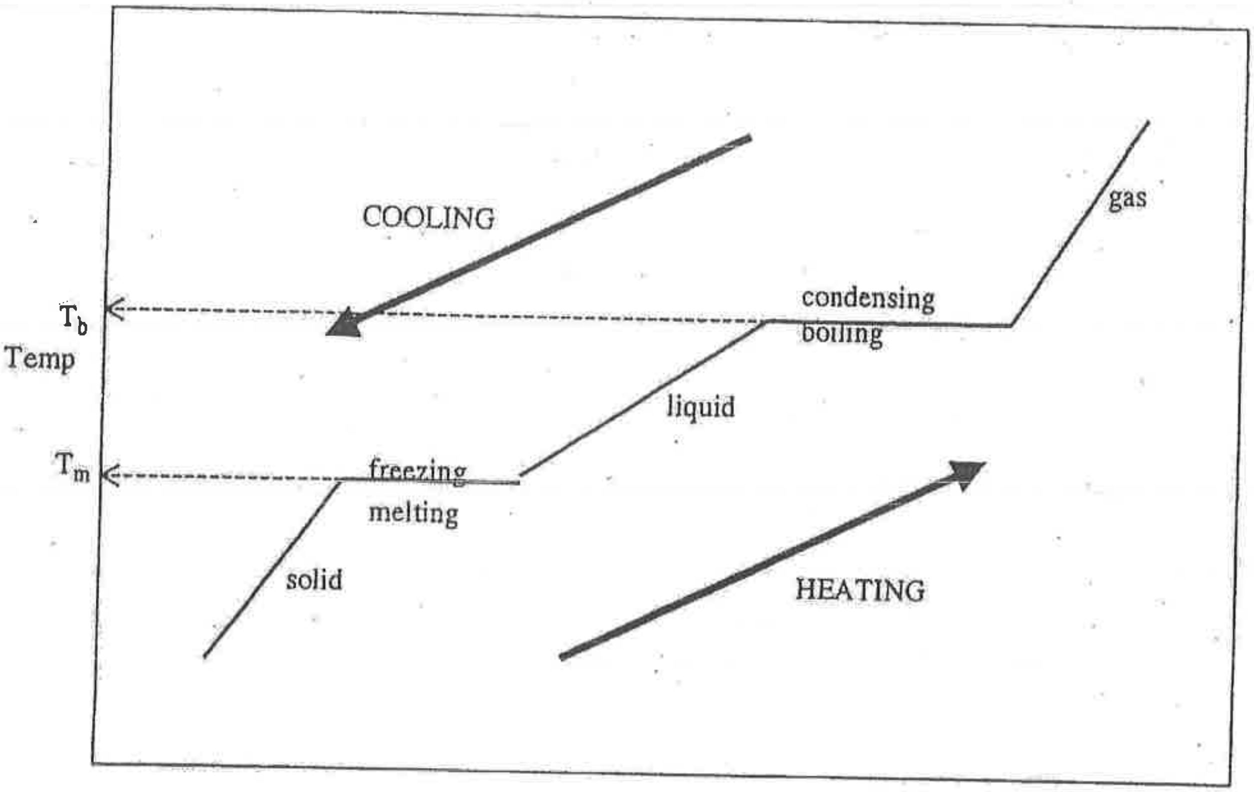
Phase Diagram



Heating Curve



Heating / Cooling Curves

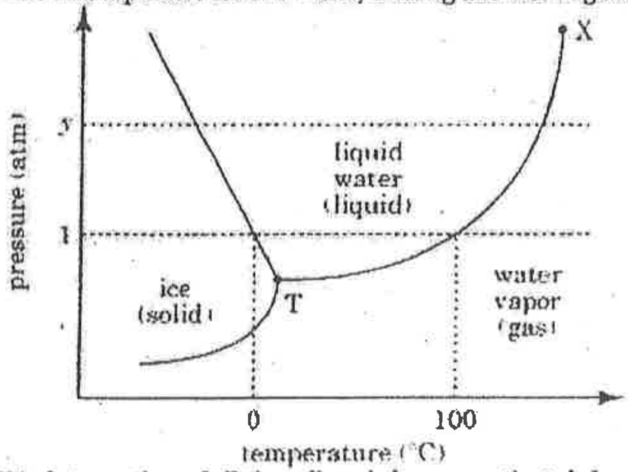


Substance	Formula	Freezing point (°C)	ΔHfus (kJ/mol)	Boiling point (°C)	ΔHvap (kJ/mol)	C - Solid (J/g°C)	C - Liquid (J/g°C)	C - Gas (J/g°C)
Acetone	C_3H_6O	-95.4	5.72	56.2	29.1		126	75.3
Ammonia	NH_3	-77.9	5.65	-33.5	23.4			35.7
Argon	Ar	-189.4	1.2	-185.9	6.5			0.519
Benzene	C_6H_6	5.5	9.87	80.1	30.8		136	81.6
Ethanol	C_2H_5OH	-114.5	4.60	78.3	43.5		2.438	65.7
Helium	He	-269.7	0.02	-268.98	0.08			5.19
Hydrogen	H_2	-259.2	0.12	-252.9	0.90			14.3
Methane	CH_4	-182.5	0.94	161.5	8.2			2.20
Methanol	CH_3OH	-97.7	3.16	64.0	35.3		81.6	43.9
Neon	Ne	-248.7	0.33	-246.1	1.76			1.03
Nitrogen	N_2	-209.9	0.72	-195.8	5.58			1.04
Oxygen	O_2	-218.4	0.44	-183.0	6.82			0.916
Water	H_2O	0.0	6.01	100.0	40.7	2.1	4.184	1.7
Aluminum	Al	660.37	10.7	2467	293	0.904		0.793
Iron	Fe	1535	14.9	2750	347	0.449		0.460
Silver	Ag	961.93	11.3	2212	255	0.235		0.193
Mercury	Hg	-38.87	2.33	356.58	59.6		0.139	0.0980

Phase Changes

der for a substance to move between the states of matter; for example, to turn from a solid into a liquid, which ... called **fusion**, or from a gas to a liquid (**vaporization**), energy must be gained or lost. The **heat of fusion** (symbolized H_{fus}) of a substance is the amount of energy that must be put into the substance for it to melt. For example, the heat of fusion of water is 6.01 kJ/mol, or in other terms, 80 cal/g. The **heat of vaporization**, not surprisingly, is the amount of energy needed to cause the transition from liquid to gas, and it is symbolized H_{vap}. You will not be required to memorize heat of fusion or vaporization values for the exam.

Changes in the states of matter are often shown on phase diagrams. Let's start with the phase diagram for water. The **phase diagram** for water is a graph of pressure versus temperature. Each of the lines on the graph represents an equilibrium position, at which the substance is present in two states at once. For example, anywhere along the line that separates ice and water, melting and freezing are occurring simultaneously.



The intersection of all three lines is known as the **triple point** (represented by a dot and a *T* on the figure). At this point, all three phases of matter are in equilibrium with each other. Point *X* represents the **critical point**, and at the critical point and beyond, the substance is forever in the vapor phase.

This diagram allows us to explain strange phenomena, such as why water boils at a lower temperature at higher altitudes, for example. At higher altitudes, the air pressure is lower, and this means that water can reach the boiling point at a lower temperature. Interestingly enough, water would boil at room temperature if the pressure was low enough!

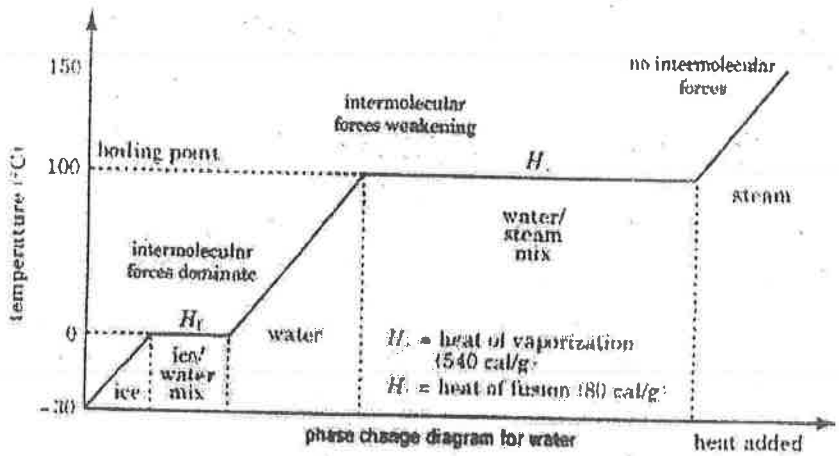
One final note: If we put a liquid into a closed container, the evaporation of the liquid will cause an initial increase in the total pressure of the system, and then the pressure of the system will become a constant. The value of this final pressure is unique to each liquid and is known as the liquid's **vapor pressure**. Water has a relatively low vapor pressure because it takes a lot of energy to break the hydrogen bonds so that molecules enter the gas phase. Water and other liquids that have low vapor pressures are said to be **nonvolatile**. Substances like rubbing alcohol and gasoline, which have relatively high vapor pressures, are said to be **volatile**.

Example

What happens to water when the pressure remains constant at 1 atm but the temperature changes from -10°C to 75°C?

Explanation

Looking at the phase change diagram for water and following the dashed line at 1 atm, you can see that water would begin as a solid (ice) and melt at 0°C. All of the water would be in liquid form by the time the temperature reached 75°C.



This is a graph of the change in temperature of a substance as energy is added in the form of heat. The pressure of the system is assumed to be held constant, at normal pressure (1 atm). As you can see from the graph below, at normal pressure water freezes at 0°C and boils at 100°C.

The plateaus on this diagram represent the points where water is being converted from one phase to another; at these stages the temperature remains constant since all the heat energy added is being used to break the attractions between the water molecules.

Specific Heat

Let's calculate the amount of energy needed to take a particular substance through a phase change. You would need to use the following equation:

$$\text{energy (in calories)} = mC_p\Delta T$$

where m = the mass of the substance (in grams)
C_p = the specific heat of the substance (in cal/g °C)
ΔT = the change in temperature of the substance (in either Kelvins or °C, but make sure all your units are compatible!)
As you can see, this requires that you know the specific heat of the substance. A substance's **specific heat** refers to the heat required to raise the temperature of 1 g of a substance by 1°C.

Work through the example below to get a feel for how to use this equation.

Example

If you had a 10.0 g piece of ice at -10.0°C, under constant pressure of 1 atm, how much energy would be needed to melt this ice and raise the temperature to 25.0°C?

Explanation

First, the temperature of the ice would need to be raised from -10.0°C to 0.00°C. This would require the following calculation. The specific heat for ice is 0.485 cal/g °C. Substituting in the formula

$$\text{Energy(cal)} = mC_p\Delta T$$

$$\text{Energy(cal)} = (10.0\text{ g})\ (0.485\text{ cal/g °C})\ (10.0\text{°C}) = 48.5\text{ cal}$$

So 48.5 calories are needed to raise temperature.
Next, we must calculate the heat of fusion of this ice: we must determine how much energy is needed to completely melt the 10 g of it.

$$\text{energy} = mH_{fus}$$

$$\text{energy} = (10.0\text{ g})\ (80\text{ cal/g}) = 800.\text{ cal}$$

So 800. cal of energy are needed to completely melt this sample of ice.
Next, we need to see how much energy would be needed to raise the temperature of water from 0.00°C to 25.0°C. The specific heat for liquid water is 1.00 cal/g °C. So again use

$$\text{energy} = mC_p\Delta T \text{ to get energy} = (10.0\text{ g})\ (1.00\text{ cal/g °C})\ (25.0\text{°C}) = 250.\text{ cal}$$

Finally, add together all of the energies to get the total: 48.5 + 800. + 250. = 1100 calories are needed to convert the ice to water at these given temperatures.

Name____Period_____________

Specific Heat Problems

Specific Heat of Water	J/g°C
Solid	2.06
Liquid	4.18
Gas	2.07

Heat of Fusion of water = 334 J/g
Heat of Vaporization of water = 2260 J/g

1. What is the energy change in Joules when 250. grams of ice at 10°C is turned into steam at 100.°C? Draw a graph to show the changes in temperature and phase changes.

2. What is the energy lost in joules when 35 grams of steam at 135°C is cooled to ice at – 3° C?

3. How many grams of ice are required to bring the temperature of 300 grams of water at 60°C to 30°C?

4. What is the final temperature of the water if 50 grams of ice are added to 500 grams of water at 40°C?

5. How much energy is required to heat 400 g of water from 20°C to 30°C?

6. As 500 g of steam are cooled and condensed from a temperature of 230°C to a liquid at 50°C, how much energy is released?

7. What amount of heat is removed from 225 g of water as it is cooled from 80°C to –80°C?

Name ______________________________ Date______________________

Specific Heat Problems

1. How much energy is required to heat 400. g of water from 20.0°C to 30.0°C?

2. As 500.g of steam is cooled and condensed froma temperature of 230.°C to a liquid at 50.0°C, how much energy is released?

3. What amount of heat is required to turn 90.0 g of water at 25.0°C into steam?

4. One hundred grams of ice at 100.K is warmed to 300.K. What amount of energy was required?

5. What amount of heat is removed from 225g of water as it is cooled from 80.0°C to -80.0°C?

6. Make an energy diagram of time vs. temperature. For the curve show steam being cooled to a liquid and then to a solid. Label the appropriate areas with the specific heats for the phases and the H_f and Hv. Include the numeric values with the labels.

Name ______________________ Date ______________ Class ______________

CHAPTER 10 REVIEW

States of Matter

SECTION 1

SHORT ANSWER Answer the following questions in the space provided.

1. Identify whether the descriptions below describe an ideal gas or a real gas.

_______________ a. The gas will not condense because the molecules do not attract each other.

_______________ b. Collisions between molecules are perfectly elastic.

_______________ c. Gas particles passing close to one another exert an attraction on each other.

2. The formula for kinetic energy is $KE = \frac{1}{2}mv^2$.

 a. As long as temperature is constant, what happens to the kinetic energy of the colliding particles during an elastic collision?

 ____

 b. If two gases have the same temperature and share the same energy but have different molecular masses, which molecules will have the greater speed?

 ____

3. Use the kinetic-molecular theory to explain each of the following phenomena:

 a. A strong-smelling gas released from a container in the middle of a room is soon detected in all areas of that room.

 ____

 b. As a gas is heated, its rate of effusion through a small hole increases if all other factors remain constant.

 ____

 ____

4. a. _______________ List the following gases in order of rate of effusion, from lowest to highest. (Assume all gases are at the same temperature and pressure.)

 (a) He (b) Xe (c) HCl (d) Cl_2

Name ____________________ Date ____________ Class ____________

SECTION 1 continued

b. Explain why you put the gases in the order above. Refer to the kinetic-molecular theory to support your explanation.

____

____

____

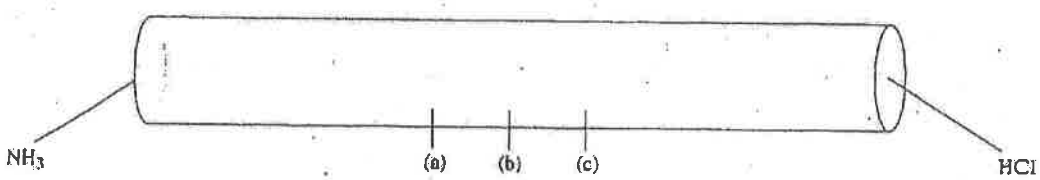
5. Explain why polar gas molecules experience larger deviations from ideal behavior than nonpolar molecules when all other factors (mass, temperature, etc) are held constant.

____

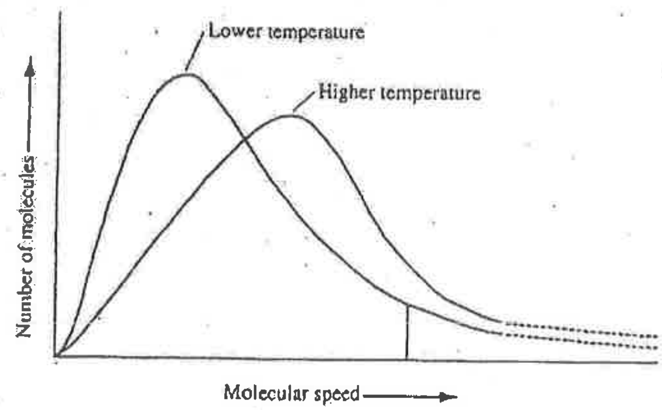
____

____

6. ______ The two gases in the figure below are simultaneously injected into opposite ends of the tube. The ends are then sealed. They should just begin to mix closest to which labeled point?



7. Explain the difference in the speed-distribution curves of a gas at the two temperatures shown in the figure below.



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Name ______________________ Date ____________ Class ____________

CHAPTER 10 REVIEW

States of Matter

SECTION 2

SHORT ANSWER Answer the following questions in the space provided.

1. ______ Liquids possess all the following properties *except*

   (a) relatively low density.
   (b) the ability to diffuse.
   (c) relative incompressibility.
   (d) the ability to change to a gas.

2. a. Chemists distinguish between intermolecular and intramolecular forces. Explain the difference between these two types of forces.

   ____

   ____

   Classify each of the following as intramolecular or intermolecular:

   ______________ b. hydrogen bonding in liquid water

   ______________ c. the O—H covalent bond in methanol, CH_3OH

   ______________ d. the bonds that cause gaseous Cl_2 to become a liquid when cooled

3. Explain the following properties of liquids by describing what is occurring at the molecular level.

   a. A liquid takes the shape of its container but does not expand to fill its volume.

   ____

   ____

   ____

   ____

   b. Polar liquids are slower to evaporate than nonpolar liquids.

   ____

   ____

   ____

Name ______________________ Date ____________ Class ____________

SECTION 2 continued

4. Explain briefly why liquids tend to form spherical droplets, decreasing surface area to the smallest size possible.

____

____

____

____

5. Is freezing a chemical change or a physical change? Briefly explain your answer.

____

____

____

____

6. Is evaporation a chemical or physical change? Briefly explain your answer.

____

____

____

____

7. What is the relationship between vaporization and evaporation?

____

____

____

____

Name ______________________ Date ____________ Class ____________

CHAPTER 10 REVIEW

States of Matter

SECTION 3

SHORT ANSWER Answer the following questions in the space provided.

1. Match description on the right to the correct crystal type on the left.

______ ionic crystal	(a) has mobile electrons in the crystal
______ covalent molecular crystal	(b) is hard, brittle, and nonconducting
______ metallic crystal	(c) typically has the lowest melting point of the four crystal types
______ covalent network crystal	(d) has strong covalent bonds between neighboring atoms

2. For each of the four types of solids, give a specific example other than one listed in Table 1 on page 340 of the text.

____

____

____

____

____

____

3. A chunk of solid lead is dropped into a pool of molten lead. The chunk sinks to the bottom of the pool. What does this tell you about the density of the solid lead compared with the density of the molten lead?

____

4. Answer *amorphous solid* or *crystalline solid* to the following questions:

______________ a. Which is less compressible?

______________ b. Which has a more clearly defined shape?

______________ c. Which is sometimes described as a supercooled liquid?

______________ d. Which has a less clearly defined melting point?

5. Explain the following properties of solids by describing what is occurring at the atomic level.

 a. Metallic solids conduct electricity well, but covalent network solids do not.

 b. The volume of a solid changes only slightly with a change in temperature or pressure.

 c. Amorphous solids do not have a definite melting point.

 d. Ionic crystals are much more brittle than covalent molecular crystals.

6. Experiments show that it takes 6.0 kJ of energy to melt 1 mol of water ice at its melting point but only about 1.1 kJ to melt 1 mol of methane, CH_4, at its melting point. Explain in terms of intermolecular forces why it takes so much less energy to melt the methane.

Name ______________________ Date ____________ Class ____________

CHAPTER 10 REVIEW

States of Matter

SECTION 4

SHORT ANSWER Answer the following questions in the space provided.

1. ______ When a substance in a closed system undergoes a phase change and the system reaches equilibrium,
   (a) the two opposing changes occur at equal rates.
   (b) there are no more phase changes.
   (c) one phase change predominates.
   (d) the amount of substance in the two phases changes.

2. Match the following definitions on the right with the words on the left.

______ equilibrium	(a) melting
______ volatile	(b) opposing changes occurring at equal rates in a closed system
______ fusion	(c) readily evaporated
______ deposition	(d) a change directly from a gas to a solid

3. Match the process on the right with the change of state on the left.

______ solid to gas	(a) melting
______ liquid to gas	(b) condensation
______ gas to liquid	(c) sublimation
______ solid to liquid	(d) vaporization

4. Refer to the phase diagram for water in Figure 16 on page 347 of the text to answer the following questions:

______________ a. What point represents the conditions under which all three phases can coexist?

______________ b. What point represents a temperature above which only the vapor phase exists?

______________ c. Based on the diagram, as the pressure on the water system increases, what happens to the melting point of ice?

d. What happens when water is at point A on the curve and the temperature increases while the pressure is held constant?

____

5. Use this general equilibrium equation to answer the following questions:

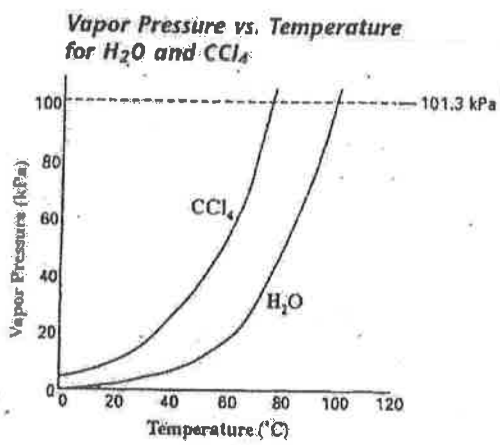
reactants ⇄ products + energy

______________ a. If the forward reaction is favored, will the concentration of reactants increase, decrease, or stay the same?

______________ b. If extra product is introduced, which reaction will be favored?

______________ c. If the temperature of the system decreases, which reaction will be favored?

6. Refer to the graph below to answer the following questions:



______________ a. What is the normal boiling point of CCl_4?

______________ b. What would be the boiling point of water if the air pressure over the liquid were reduced to 60 kPa?

______________ c. What must the air pressure over CCl_4 be for it to boil at 50°C?

d. Although water has a lower molar mass than CCl_4, it has a lower vapor pressure when measured at the same temperature. What makes water vapor less volatile than CCl_4?

Name ______________________ Date ______________ Class ____________

CHAPTER 10 REVIEW

States of Matter

SECTION 5

SHORT ANSWER Answer the following questions in the space provided.

1. Indicate whether each of the following is a *physical* or *chemical* property of water.

______________ a. The density of ice is less than the density of liquid water.

______________ b. A water molecule contains one atom of oxygen and two atoms of hydrogen.

______________ c. There are strong hydrogen bonds between water molecules.

______________ d. Ice consists of water molecules in a hexagonal arrangement.

2. Compare a polar water molecule with a less-polar molecule, such as formaldehyde, CH_2O. Both are liquids at room temperature and 1 atm pressure.

______________ a. Which liquid should have the higher boiling point?

______________ b. Which liquid is more volatile?

______________ c. Which liquid has a higher surface tension?

______________ d. In which liquid is NaCl, an ionic crystal, likely to be more soluble?

3. Describe hydrogen bonding as it occurs in water in terms of the location of the bond, the particles involved, the strength of the bond, and the effects this type of bonding has on physical properties.

____

____

____

____

____

____

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Name ______________________ Date ____________ Class ____________

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

4. The molar enthalpy of vaporization of water is 40.79 kJ/mol, and the molar enthalpy of fusion of ice is 6.009 kJ/mol. The molar mass of water is 18.02 g/mol.

______________ a. How much energy is absorbed when 30.3 g of liquid water boils?

______________ b. An energy unit often encountered is the calorie (4.18 J = 1 calorie). Determine the molar enthalpy of fusion of ice in calories per gram.

5. A typical ice cube has a volume of about 16.0 cm^3. Calculate the amount of energy needed to melt the ice cube. (Density of ice at 0.°C = 0.917 g/mL; molar enthalpy of fusion of ice = 6.009 kJ/mol; molar mass of H_2O = 18.02 g/mol.)

______________ a. Determine the mass of the ice cube.

______________ b. Determine the number of moles of H_2O present in the sample.

______________ c. Determine the number of kilojoules of energy needed to melt the ice cube.

Name ______________________ Date ______________ Class ______________

CHAPTER 10 REVIEW

States of Matter

MIXED REVIEW

SHORT ANSWER Answer the following questions in the space provided.

1. ______ The average speed of a gas molecule is most directly related to the

(a) polarity of the molecule.
(b) pressure of the gas.
(c) temperature of the gas.
(d) number of moles in the sample.

2. Use the kinetic-molecular theory to explain the following phenomena:

a. When 1 mol of a real gas is condensed to a liquid, the volume shrinks by a factor of about 1000.

____

____

____

b. When a gas in a rigid container is warmed, the pressure on the walls of the container increases.

____

____

____

____

3. ______ Which of the following statements about liquids and gases is *not* true?

(a) Molecules in a liquid are much more closely packed than molecules in a gas.
(b) Molecules in a liquid can vibrate and rotate, but they are bound in fixed positions.
(c) Liquids are much more difficult to compress into a smaller volume than are gases.
(d) Liquids diffuse more slowly than gases.

4. Answer *solid* or *liquid* to the following questions:

______________ a. Which is less compressible?

______________ b. Which is quicker to diffuse into neighboring media?

______________ c. Which has a definite volume and shape?

______________ d. Which has molecules that are rotating or vibrating primarily in place?

Name ______________________ Date ____________ Class ____________

MIXED REVIEW continued

5. Explain why almost all solids are denser than their liquid states by describing what is occurring at the molecular level.

____

____

6. A general equilibrium equation for boiling is

$$\text{liquid} + \text{energy} \rightleftarrows \text{vapor}$$

Indicate whether the forward or reverse reaction is favored in each of the following cases:

____________ a. The temperature of the system is increased.

____________ b. More molecules of the vapor are added to the system.

____________ c. The pressure on the system is increased.

7. ____________ Freon-11, CCl_3F has been commonly used in air conditioners. It has a molar mass of 137.35 g/mol and its enthalpy of vaporization is 24.8 kJ/mol at its normal boiling point of 24°C. Ideally how much energy in the form of heat is removed from a room by an air conditioner that evaporates 1.00 kg of freon-11?

8. Use the data table below to answer the following:

Composition	Molar mass (g/mol)	Enthalpy vaporization (kJ/mol)	Normal boiling point (°C)	Critical temperature (°C)
He	4	0.08	−269	−268
Ne	20	1.8	−246	−229
Ar	40	6.5	−186	−122
Xe	131	12.6	−107	+17
H_2O	18	40.8	+100	+374
HF	20	25.2	+20	+188
CH_4	16	8.9	−161	−82
C_2H_6	30	15.7	−89	+32

____________ a. Among *nonpolar* liquids, those with higher molar masses tend to have normal boiling points that are (higher, lower, or about the same).

____________ b. Among compounds of approximately the same molar mass, those with greater polarities tend to have enthalpies of vaporization that are (higher, lower, or about the same).

c. Which is the only noble gas listed that is stable as a liquid at 0°C? Explain your answer using the concept of critical temperature.

____

____

Chapter 11 Gas Laws

Ideal Gas Law Equation

$$PV = nRT$$

Variable	Name	Definition	Units	Measured by:
P	Pressure	The force per unit area that the gas exerts on the any surface.	atm, mmHg, kPa	Barometer, pressure gauge
V	Volume	The amount of space occupied.	L	Graduated Cylinder
T	Temperature	The measure of the average kinetic energy in a system	K	Thermometer
R	Ideal Gas Law Constant	$0.0821 \frac{L \cdot atm}{mol \cdot K}$ $8.31 \frac{L \cdot kPa}{mol \cdot K}$		Calculated
N	Moles	6.022×10^{23} Particles	Mol	Number of particles in 12 grams of carbon-12.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

P = Pressure of the gas
T = Temperature of the gas
V = Volume of the gas

Combined Gas Law: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Boyle's Law: $P_1V_1 = P_2V_2$

Charles' Law: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Name ______________________ Period ________

CHEMISTRY GAS LAW'S WORKSHEET

Boyle's Law	Charles' Law	Guy-Lassac's Law	Combined Gas Law
For a given mass of gas at constant temperature, the volume of a gas varies inversely with pressure	The volume of a fixed mass of gas is directly proportional to its Kelvin temperature if the pressure is kept constant.	The pressure of a gas is directly proportional to the Kelvin temperature if the volume is kept constant.	Combines Boyle's, Charles', and the Temperature-Pressure relationship into one equation. Each of these laws can be derived from this law.
$PV = k$ $P_1V_1 = P_2V_2$	$\frac{V}{T} = k$ $V_1T_2 = V_2T_1$ $\frac{V_1}{T_1} = \frac{V_2}{T_2}$	$\frac{P}{T} = k$ $P_1T_2 = P_2T_1$ $\frac{P_1}{T_1} = \frac{P_2}{T_2}$	$\frac{PV}{T} = k$ $V_1P_1T_2 = V_2P_2T_1$ $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Dalton's Law	Ideal Gas Law	Graham's Law
At constant volume and temperature, the total pressure exerted by a mixture of gases is equal to the sum of the pressures exerted by each gas,	The Ideal Gas Law relates the pressure, temperature, volume, and mass of a gas through the gas constant "R".	The rate of effusion/diffusion of two gases (A and B) are inversely proportional to the square roots of their formula masses. [*It can be a ratio of molecular speeds, effusion /diffusion times, distance traveled by molecules, or amount of gas effused*]
$P_{total} = P_1 + P_2 + P_3 + \ldots P_n$	$PV = nRT$	$\frac{Rate_A}{Rate_B} = \frac{\sqrt{molar\ mass_B}}{\sqrt{molar\ mass_A}}$

Abbreviations	Standard Conditions
atm = atmosphere mm Hg = millimeters of mercury torr = another name for mm Hg Pa = Pascal kPa = kilopascal K = Kelvin °C = degrees Celsius	0°C = 273 K 1.00 atm = 760.0 mm Hg = 76 cm Hg =101.325 kPa = 101, 325 Pa = 29.9 in Hg
Conversions	**Gas Law's Equation Symbols**
K = °C + 273 $F^{\circ} = 1.8C^{\circ} + 32$ $C^{\circ} = \frac{F^{\circ} - 32}{1.8}$ 1 cm^3 (cubic centimeter) = 1 mL (milliliter) 1 dm^3 (cubic decimeter) = 1 L (liter) = 1000 mL	Subscript (1) = old condition or initial condition Subscript (2) = new condition or final condition Temperature must be in Kelvins n = number of moles = grams/Molar mass R = 8.31 L-kPa/ mol-K = 0.0821 L-atm/mol-K = 62.4 L-Torr/mol-K You must have a common set of units in the problem

Name ____ Date__________

Chapter 11: Gases

Pressure is the force per unit area on a surface. SI unit of force is the **newton.**

Measuring Pressure:

Barometer:

Pressure conversions: 760mm Hg = 1 atm = = 101.3kPa = 101300 Pa = 760 torr = ______ lb/in^2

Manometer:

Dalton's Law of Partial Pressures: Total pressure of a gas mixture is the sum of the partial pressures of the component gases.

Collecting gases by water displacement: $P_{atm} = P_{gas} + P_{H2O}$

- Gas Laws: Boyle's, Charles', Gay-Lussac, Combined, Ideal
- Avogadro's Law: Equal volumes of gases at the same temperature and pressure have the same number of molecules. Gas volume is directly proportional to the amount of gas.
- Molar volume of a gas. Any gas will occupy a volume of 22.4L if it is at STP conditions.
- Graham's Law of Effusion:

Name ______________________ Date ____________ Class __________

13–2 Practice Problems

1. The air pressure for a certain tire is 109 kPa. What is this pressure in atmospheres?

2. The air pressure inside a submarine is 0.62 atm. What would be the height of a column of mercury balanced by this pressure?

3. The weather news gives the atmospheric pressure as 1.07 atm. What is this atmospheric pressure in mm Hg?

4. An experiment at Sandia National Labs in New Mexico is performed at an atmospheric pressure of 758.7 mm Hg. What is this pressure in atm?

5. A bag of potato chips is sealed in a factory near sea level. The atmospheric pressure at the factory is 761.3 mm Hg. The pressure inside the bag is the same. What is the pressure inside the bag of potato chips in Pa?

6. The same bag of potato chips from Problem 5 is shipped to a town in Colorado, where the atmospheric pressure is 99.82 kPa. What is the difference (in Pa) between the pressure in the bag and the atmospheric pressure of the town?

7. The pressure gauge on a compressed air tank reads 43.2 lb/in^2. What is the pressure in atm?

8. The pressure in the tire of an automobile is 34.8 lb/in^2. What is the pressure in kPa?

9. A gas container is fitted with a manometer. The level of the mercury is 15 mm lower on the open side. Using a laboratory barometer, you find that atmospheric pressure is 750 mm Hg. What is the pressure, in atmospheres, of the gas in the container?

10. A soccer ball is attached to an open-ended manometer. The mercury level in the manometer is 10 mm higher on the side attached to the ball than on the side open to the atmosphere. Atmospheric pressure has already been determined to be 770 mm Hg. What is the gas pressure in the ball?

11. One end of an open-ended manometer is connected to a canister filled with a gas at a pressure of 771.0 mm Hg. The mercury level on the side open to the atmosphere is 11.2 mm higher than on the side connected to the canister. What is the atmospheric pressure in mm Hg?

12. Suppose you are measuring the pressure inside a sealed cabinet using an open-ended manometer. The atmospheric pressure is 762.4 mm Hg. If the mercury level on the side open to the atmosphere is 3.6 mm higher than on the side attached to the cabinet, what is the pressure inside the cabinet in units of kPa?

13. The U-tube of a manometer is 26.4 cm tall. With both ends open, it is filled until the mercury level in each tube is 13.2 cm from the top. What is the largest difference in pressure this manometer can measure in units of mm Hg?

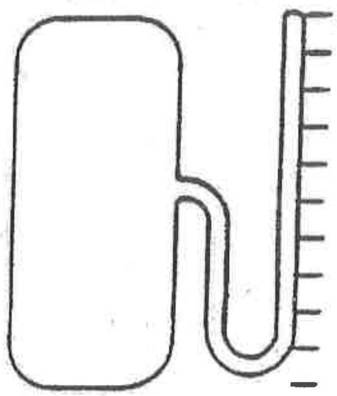
14. A manometer contains a sample of nitrogen gas at a pressure of 88.3 kPa. The level of mercury in the U-tube is 12.8 mm lower on the end open to the atmosphere. What is the atmospheric pressure in kPa?

15. One end of an open-ended manometer is connected to a canister of unknown gas. The atmospheric pressure is 1.03 atm. The mercury level is 18.6 mm higher in the U-tube on the side open to the atmosphere than on the side attached to the canister. What is the pressure of the gas in mm Hg?

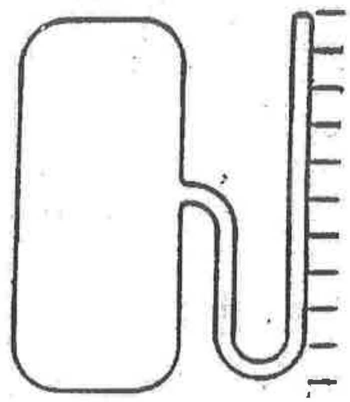
Closed-end manometers are used to find the actual pressure of a confined gas

Either calculate the pressure of the gas in the container or shade the tubes of the manometers to show the mercury levels that indicate the P_{gas}. Use the attached rulers to show the measurements. Each marking spaced at 1 cm intervals represents 100 mm of mercury.

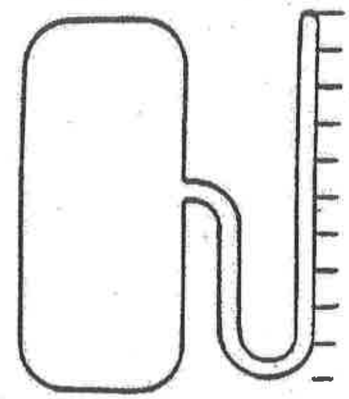
A. P_{gas} = 250 mm Hg



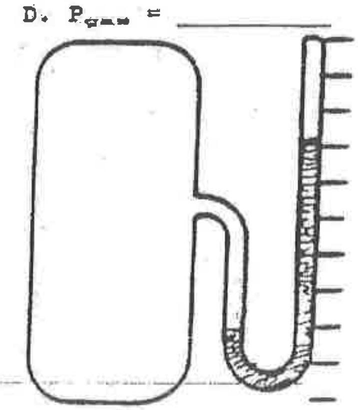
B. P_{gas} = 400 mm Hg



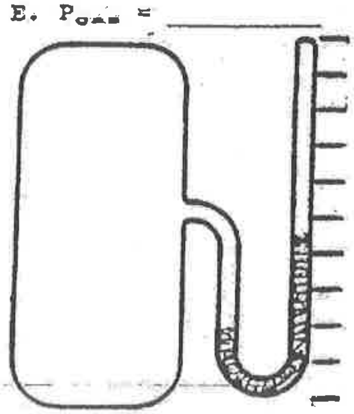
C. P_{gas} = 340 mm Hg



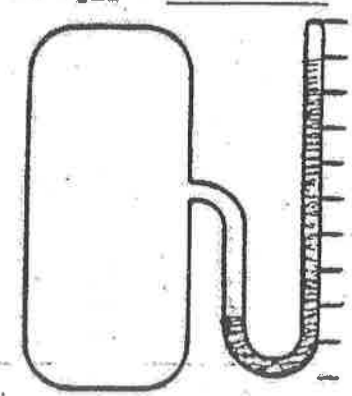
D. P_{gas} = __________



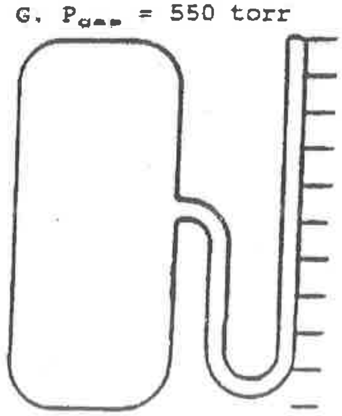
E. P_{gas} = __________



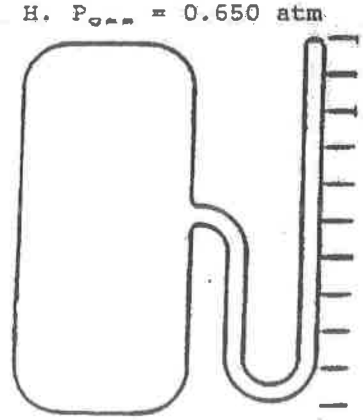
F. P_{gas} = __________



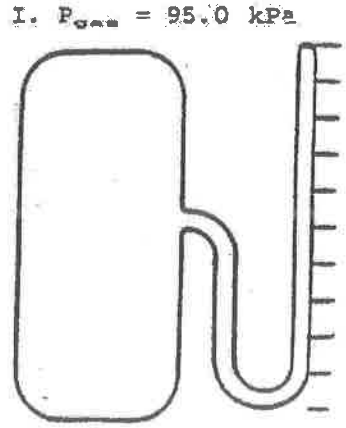
G. P_{gas} = 550 torr



H. P_{gas} = 0.650 atm



I. P_{gas} = 95.0 kPa

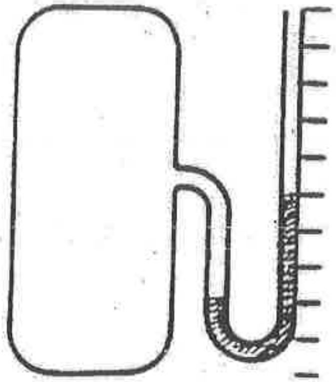


Name ______________________________ Class______ Date______________

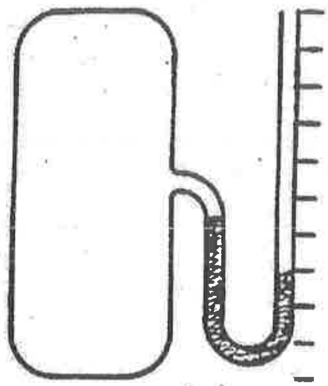
. Either calculate the pressure of the gas in the container or shade the tubes of the manometers to show the mercury levels that indicate the P_{gas} and P_{atm}. Use the attached rulers to show the measurements. Each marking spaced at 1 cm intervals represents 100 mm of mercury.

If P_{atm} = 780.0 mm Hg

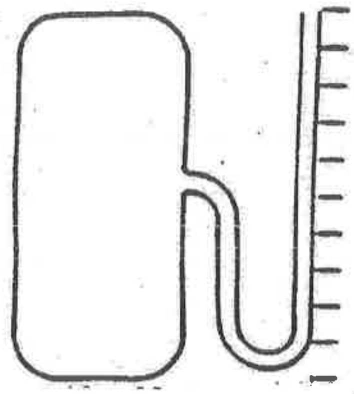
A. P_{gas} = __________



B. P_{gas} = __________

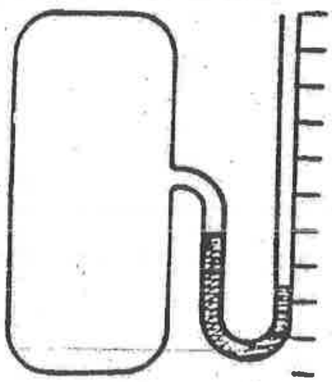


C. P_{gas} = 1000. mm Hg

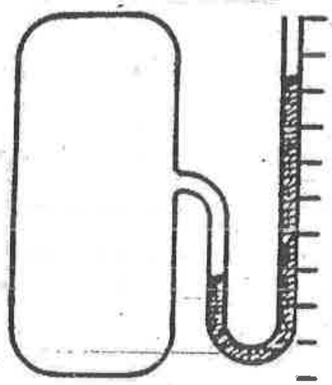


If P_{atm} = 95.0 kPa

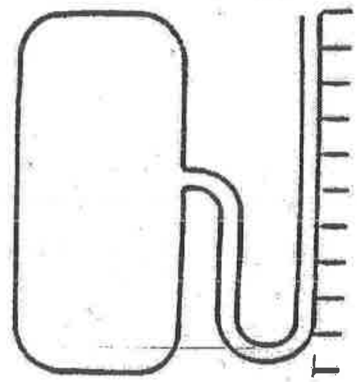
D. P_{gas} = __________



E. P_{gas} = __________

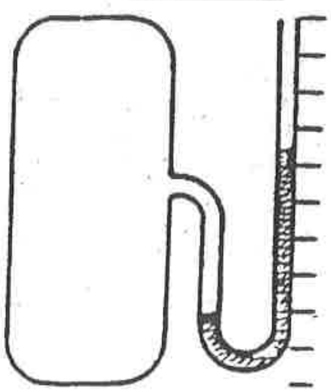


F. P_{gas} = 612 mm Hg

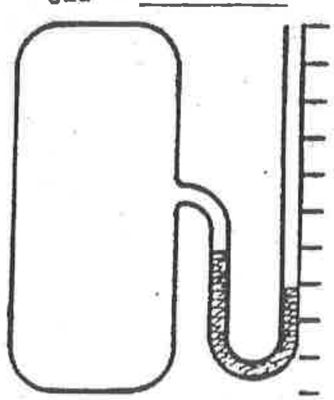


If P_{atm} = 1.10 atm

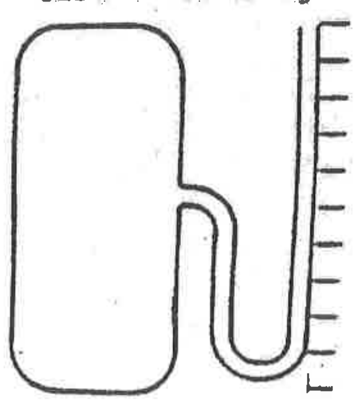
G. P_{gas} = __________



H. P_{gas} = __________



I. P_{gas} = 1400. mm Hg



Name ______________________________ **Date** ______________

Boyle's Law

Pressure-volume relationship

perature must remain constant

This law follows an inverse relationship between pressure and volume. When one is up the other is down.

Practice problems:

1. A gas occupies a volume of 458mL at a pressure of 1.01kPa and a temperature of 295K. When the pressure is changed, the volume becomes 477 mL. If there has been no change in temperature, what is the new pressure?

2. A gas occupies a volume of 2.45L at a pressure of 1.03 atm and a temperature of 293K. What volume will the gas occupy if the pressure changes to 0.980 atm and the temperature remains unchanged?

3. A discarded spray paint can contains only a small volume of the propellant gas at a pressure of 34,470 Pa. The volume of the can is 473.18mL. If the can is run over by a garbage truck and flattened to a volume of 13.16mL, what is the pressure in Pa assuming the can does not leak?

4. A sample of gas occupying a volume of 10.0L with a pressure of 78.6 lb/in2 is squished into a 2.8 L container. What is the new pressure of the gas, assuming the temperature remains constant?

Charles's Law

Temperature- volume relationship *(Pressure must remain constant)*
Temperature **MUST BE** in Kelvin.

This law follows a direct relationship between temperature and volume. When one is up the other is also up. Jacques Charles was one of three passengers in the world's second balloon ascension that carried humans. The launch of the hydrogen filled balloon took place on 12-1-1783 in Paris

Sample problems:

1. A gas occupies 25.0 L of space at 35° C. At what temperature would it occupy 5.0 L of space if the pressure remains constant?

2. The volume of a balloon is 6.2L at 298K. What will be the volume at 302K? Pressure is constant.

3. A sample of gas at 83°C occupies a volume of $1400m^3$. At what temperature will it occupy $1200m^3$?

4. What will be the volume of a gas sample at 309K if its volume at 215 K is 3.42L? Pressure is constant.

Gay –Lussac's Law

Temperature and Pressure Relationship: *Volume must remain constant.* This relationship is also direct. As the temperature of the gas goes up, so does its pressure.

Practice problems:

1 . At a temperature of 33.0°C a sample of confined gas exerts a pressure of 53.3kPa. If the volume remains constant, at what temperature will the pressure reach 133kPa?

2. A gas confined in a rigid container exerts a pressure of 33.5 kPa at a temperature of 17.0° C. What will the pressure of this gas be if it is cooled to a temperature of –23.0° C?

Combined Gas Law:

Use this law if temperature, pressure or volume is not a constant. It is the result of the combination of Boyles, ırles and Gay-Lussac's Laws.

1. A gas is confined in a 200L container at a pressure of 2.0atm and a temperature of 25 °C. If the gas is transferred to a container that is 100L and the temperature is raised to 30°C, what is the new pressure?

2. A gas is squished inside a container that is 50L. It is at a pressure of 102kPa and the temperature is 50°C. If the pressure is decreased to 90kPa and the temperature is decreased to 0°C, what size container can it now be squished in?

Name ____________________ Date ____________ Class ____________

13–4 Practice Problems

1. What volume would be occupied by 100. g of oxygen gas at a pressure of 1.50 atm and a temperature of 25°C?

2. An air-filled balloon has a volume of 225 L at 0.94 atm and 25°C. Soon after, the pressure changes to 0.99 atm and the temperature changes to 0°C. What is the new volume of the balloon?

3. A gas confined in a 515-cm^3 container exerts a pressure of 107.4 kPa at 38.6°C. At what Celsius temperature will it exert a pressure of 635.7 kPa if it is placed into a 644-cm^3 container?

4. A balloon is inflated with 0.2494 g of helium to a pressure of 1.26 atm. If the desired volume of the balloon is 1.250 L, what must the temperature be in °C?

5. A welder's acetylene tank has a volume of 75.0 L. It is stored at a temperature of 23.24°C and has a pressure of 7667 kPa. How many moles of acetylene are in the tank?

6. How many grams of argon would it take to fill a light bulb with a volume of 0.475 L at STP?

7. Dry ice is carbon dioxide in the solid state. 1.28 grams of dry ice are placed into a 5.00 L evacuated chamber that is maintained at 35.1°C. What is the pressure in the chamber in kPa after all the dry ice has sublimed into CO_2 gas?

8. A sample of Br_2 gas is loaded into an evacuated demonstration bottle at STP. The volume of the bottle is 0.25 L. How many moles of Br_2 gas will be contained in the bottle?

9. A sample of gas occupies 0.308 m^3 at a temperature of 325 K and a pressure of 149 kPa. Calculate the number of moles of the gas that are present.

10. What pressure is exerted by 0.625 mole of a gas in a 45.4 L container at −24.0°C?

Gases Review

FOR EACH OF THE FOLLOWING, NAME THE GAS LAW AND SOLVE THE PROBLEM:

1. Hydrogen is collected over water at 0.975 atm and 28°C. What is the partial pressure of H_2?
2. How many moles of chloroform, $CHCl_3$, are required to fill a 253-mL flask at 100.0°C and 940 torr?
3. You want the pressure inside a bottle to be 75.0 kPa at 23°C. At what temperature in Celsius should you seal the bottle when the pressure is 1.12 atm?
4. A diver's lungs hold about 20.0 L of air underwater at a pressure of 875 mm Hg. Assuming he holds his breath and his lungs don't burst, what will be the volume of air in his lungs at standard pressure on the water's surface.
5. A soccer ball containing 12.0 dm^3 of gas at 21°C is left outside on a cold, winter day. What is the temperature outside in Celsius if the ball shrunk to 10,500 cm^3?
6. What pressure is required to compress a gas that occupies 6500 L at 25°C and 1.0 atm to a volume of 40.0 L at 18°C?
7. Oxygen gas diffuses how many times faster than sulfur dioxide?
8. When a canning jar is sealed at 100°C the pressure inside is 101.3 kPa. What is the pressure inside the jar when it cools to room temperature, about 21°C?
9. CO_2 gas is collected over water at 100.3 kPa and 19°C. Find the pressure of the dry gas.
10. A Marshmallow Peep® has a volume of about 45.0 cm^3 at 101 kPa. What pressure is required to increase its size to 150.0 cm^3 assuming no air escapes from the Peep®.
11. What is the temperature of a 0.00893 mol sample of neon gas that has a volume of 302 mL and a pressure of 0.941 atm?
12. A gas occupies 4.78 L at 78.1 kPa and 25°C. What will the volume be at 0.975 atm and 15°C?
13. What is the molar mass of a gas that diffuses 0.71 times as fast as oxygen?
14. A shampoo bottle contains 443 mL of air at 65°C. What is its volume when it cools to 22°C?
15. A balloon is filled with helium to a volume of 12.5 liters at 25°C and 101 kPa. How many grams of helium are in the balloon?
16. A sample of propane has a volume of 250.0 L at 125 kPa and 38°C. What volume will this sample have at 100.0 kPa and 95°C?
17. At a certain temperature and pressure, chlorine molecules have an average speed of 0.0380 m/s. What is the average speed of SO_2 molecules under the same conditions?
18. The pressure in a can of hairspray is 2.50 atm at 298 K. What is the pressure in the can when it is heated to 398 K?
19. What is the molar mass of an unknown gas if it diffuses 0.906 times as fast as argon?

SOLVE THE FOLLOWING GAS STOICHIOMETRY PROBLEMS:

20. What volume of chlorine is required to produce 25.4 g of copper(II) chloride at 18°C and 2.13 atm?
$$Cu + Cl_2 \rightarrow CuCl_2$$
21. At 778 mm Hg and 25°C, how many grams of zinc are required to produce 25.2 liters of hydrogen gas?
$$\_\_Zn + \_\_HCl \rightarrow \_\_ZnCl_2 + \_\_H_2$$
22. If 5.45 g of potassium chlorate decompose, how many liters of oxygen gas are given off at 1.58 atm and 32°C?
$$2KClO_3 \rightarrow 2KCl + 3O_2$$
23. When aluminum is burned in 15.0 L of oxygen at 97.3 kPa and 21°C, how many grams of aluminum oxide are formed?
$$4Al + 3O_2 \rightarrow 2Al_2O_3$$
24. If 12.8 g of $CaCO_3$ decomposes at 38°C and 0.96 atm, how many dm^3 of CO_2 are formed in addition to CaO?
$$CaCO_3 \rightarrow CaO + CO_2$$
25. What mass of glucose ($C_6H_{12}O_6$) is required to produce 150 cm^3 of carbon dioxide at 102 kPa and 23°C?
$$C_6H_{12}O_6 + 2O_2 \rightarrow 2CH_3COOH + 2CO_2 + 2H_2O$$

Name ________________ Date __________ Class __________

CHAPTER 11 REVIEW

Gases

SECTION 1

SHORT ANSWER Answer the following questions in the space provided.

1. ______ *Pressure* = $\frac{\textit{force}}{\textit{surface area}}$. For a constant force, when the surface area is tripled the pressure is

(a) doubled.
(b) a third as much.
(c) tripled.
(d) unchanged.

2. ________________ Rank the following pressures in increasing order.

(a) 50 kPa   (c) 76 torr
(b) 2 atm   (d) 100 N/m^2

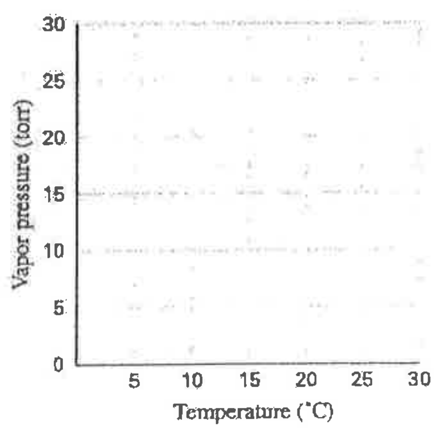
3. Explain how to calculate the partial pressure of a dry gas that is collected over water when the total pressure is atmospheric pressure.

____

____

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

4. a. Use five to six data points from **Appendix Table A-8** in the text to sketch the curve for water vapor's partial pressure versus temperature on the graph provided below.



________________ **b.** Do the data points lie on a straight line?

________________ **c.** Based on your sketch, predict the approximate partial pressure for water at 11°C.

5. Convert a pressure of 0.200 atm to the following units:

______________ **a.** mm Hg

______________ **b.** kPa

6. When an explosive like TNT is detonated, a mixture of gases at high temperature is created. Suppose that gas X has a pressure of 50 atm, gas Y has a pressure of 20 atm, and gas Z has a pressure of 10 atm.

______________ **a.** What is the total pressure in this system?

______________ **b.** What is the total pressure in this system in kPa?

7. The height of the mercury in a barometer is directly proportional to the pressure on the mercury's surface. At sea level, pressure averages 1.0 atm and the level of mercury in the barometer is 760 mm (30. in.). In a hurricane, the barometric reading may fall to as low as 28 in.

______________ **a.** Convert a pressure reading of 28 in. to atmospheres.

______________ **b.** What is the barometer reading, in mm Hg, at a pressure of 0.50 atm?

c. Can a barometer be used as an altimeter (a device for measuring altitude above sea level)? Explain your answer.

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Name ______________________ Date ______________ Class ______________

CHAPTER 11 REVIEW

Gases

SECTION 2

SHORT ANSWER Answer the following questions in the space provided.

1. State whether the pressure of a fixed mass of gas will increase, decrease, or stay the same in the following circumstances:

    ______________ **a.** temperature increases, volume stays the same

    ______________ **b.** volume increases, temperature stays the same

    ______________ **c.** temperature decreases, volume stays the same

    ______________ **d.** volume decreases, temperature stays the same

2. Two sealed flasks, A and B, contain two different gases of equal volume at the same temperature and pressure. Assume that flask A is warmed as flask B is cooled. Will the pressure in the two flasks remain equal? If not, which flask will have the higher pressure?

____

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PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

3. A bicycle tire is inflated to 55 lb/in.^{2} at 15°C. Assume that the volume of the tire does not change appreciably once it is inflated.

    a. If the tire and the air inside it are heated to 30°C by road friction, does the pressure in the tire increase or decrease? (Assume the volume of air in the tire remains constant.)

    ____

    ____

    b. Because the temperature has doubled, does the pressure double to 110 psi?

    ____

    ____

    c. What will the pressure be when the temperature has doubled? Express your answer in pounds per square inch.

    ____

Name ________________________ Date ______________ Class ______________

SECTION 2 continued

4. ________________ A 24 L sample of a gas at fixed mass and constant temperature exerts a pressure of 3.0 atm. What pressure will the gas exert if the volume is changed to 16 L?

5. ________________ A common laboratory system to study Boyle's law uses a gas trapped in a syringe. The pressure in the system is changed by adding or removing identical weights on the plunger. The original gas volume is 50.0 mL when two weights are present. Predict the new gas volume when four more weights are added.

6. ________________ A sample of argon gas occupies a volume of 950 mL at 25.0°C. What volume will the gas occupy at 50.0°C if the pressure remains constant?

7. ________________ A 500.0 mL gas sample at STP is compressed to a volume of 300.0 mL and the temperature is increased to 35.0°C. What is the new pressure of the gas in pascals?

8. ________________ A sample of gas occupies 1000. mL at standard pressure. What volume will the gas occupy at a pressure of 600. mm Hg if the temperature remains constant?

Name ______________________ Date ______________ Class ______________

CHAPTER 11 REVIEW

Gases

SECTION 3

SHORT ANSWER Answer the following questions in the space provided.

1. ______ The molar mass of a gas at STP is the density of that gas

   (a) multiplied by the mass of 1 mol.
   (b) divided by the mass of 1 mol.
   (c) multiplied by 22.4 L.
   (d) divided by 22.4 L.

2. ______ For the expression $V = \frac{nRT}{P}$, which of the following will cause the volume to increase?

   (a) increasing P
   (b) decreasing T
   (c) increasing T
   (d) decreasing n

3. Two sealed flasks, A and B, contain two different gases of equal volume at the same temperature and pressure.

   ______________ a. The two flasks must contain an equal number of molecules. True or False?

   ______________ b. The two samples must have equal masses. True or False?

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

4. Use the data in the table below to answer the following questions.

Formula	Molar mass (g/mol)
N_2	28.02
CO	28.01
C_2H_2	26.04
He	4.00
Ar	39.95

(Assume all gases are at STP.)

______________ a. Which gas contains the most molecules in a 5.0 L sample?

______________ b. Which gas is the least dense?

______________ c. Which two gases have virtually the same density?

______________ d. What is the density of N_2 measured at STP?

______________ a. How many moles of methane, CH_4 are present in 5.6 L of the gas at STP?

______________ b. How many moles of gas are present in 5.6 L of any ideal gas at STP?

______________ c. What is the mass of the 5.6 L sample of CH_4?

6. ______________ a. A large cylinder of He gas, such as that used to inflate balloons, has a volume of 25.0 L at 22°C and 5.6 atm. How many moles of He are in such a cylinder?

______________ b. What is the mass of the He calculated in part a?

7. When C_3H_4 combusts at STP, 5.6 L of C_3H_4 are consumed according to the following equation:

$$C_3H_4(g) + 4O_2(g) \rightarrow 3CO_2(g) + 2H_2O(l)$$

______________ a. How many moles of C_3H_4 react?

______________ b. How many moles of O_2, CO_2, and H_2O are either consumed or produced in the above reaction?

____

____

______________ c. How many grams of C_3H_4 are consumed?

______________ d. How many liters of CO_2 are produced?

______________ e. How many grams of H_2O are produced?

Name ______________________ Date ______________ Class ______________

CHAPTER 11 REVIEW

Gases

SECTION 4

SHORT ANSWER Answer the following questions in the space provided.

1. ______________ List the following gases in order of increasing rate of effusion. (Assume all gases are at the same temperature and pressure.)

(a) He (b) Xe (c) HCl (d) Cl_2

2. Explain your reasoning for the order of gases you chose in item 1 above. Refer to the kinetic-molecular theory to support your explanation and cite Graham's law of effusion.

____

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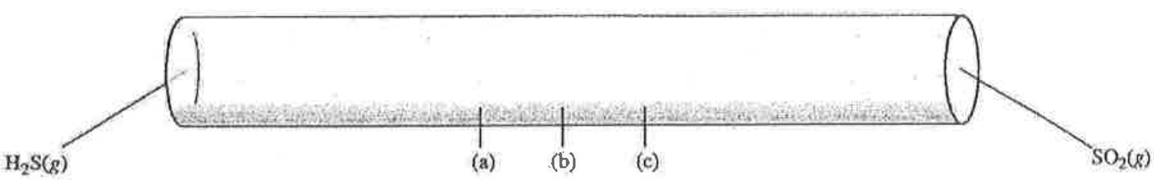
____

____

3. ______ The two gases in the figure below are simultaneously injected into opposite ends of the tube. At which labeled point should they just begin to mix?



4. State whether each example describes effusion or diffusion.

______________ **a.** As a puncture occurs, air moves out of a bicycle tire.

______________ **b.** When ammonia is spilled on the floor, the house begins to smell like ammonia.

______________ **c.** The smell of car exhaust pervades an emissions testing station.

5. Describe what happens, in terms of diffusion, when a bottle of perfume is opened.

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Name ______________________ Date ______________ Class ______________

SECTION 4 continued

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

6. ______________ a. The molar masses of He and of HCl are 4.00 g/mol and 36.46 g/mol, respectively. What is the ratio of the mass of He to the mass of HCl rounded to one decimal place?

______________ b. Use your answer in part a to calculate the ratio of the average speed of He to the average speed of HCl.

______________ c. If helium's average speed is 1200 m/s, what is the average speed of HCl?

7. ______________ An unknown gas effuses through an opening at a rate 3.16 times slower than neon gas. Estimate the molar mass of this unknown gas.