Unit 5 Interactive Notebook
Gas Laws and Kinetic Molecular Theory

Grant Union High School
January 6, 2014 – January 29, 2014

Student Mastery Scale of Learning Goals

Date	Page	Std	Learning Goal	Homework	Mastery
1/6/14					
1/7/14					
1/8/14					
1/9/14					
1/10/14					
1/13/14					
1/14/14					
1/15/14					

1/16/14					
1/21/14					
1/22/14					
1/23/14					
1/24/14					
1/27/14					
1/28/14					
1/29/14			**Unit 5 EXAM**		

California Standard Gas Laws

4. The kinetic molecular theory describes the motion of atoms and molecules and explains the properties of gases. As a basis for understanding this concept:

a. *Students know* the random motion of molecules and their collisions with a surface create the observable pressure on that surface. Fluids, gases or liquids, consist of molecules that freely move past each other in random directions. **Intermolecular forces** hold the atoms or molecules in liquids close to each other. Gases consist of tiny particles, either atoms or molecules, spaced far apart from each other and free to move at high speeds. ***Pressure*** is defined as force per unit area. The force in fluids comes from collisions of atoms or molecules with the walls of a container. Air pressure is created by the weight of the gas in the atmosphere striking surfaces. Gravity pulls air molecules toward Earth, the surface that they strike. Water pressure can be understood in the same fashion, but the pressures are much greater because of the greater density of water. Pressure in water increases with depth, and pressure in air decreases with altitude and vice versa. *However,* pressure is felt equally in all directions in fluids because of the random motion of the molecules.

4. b. *Students know* the random motion of molecules explains the diffusion of gases. Another result of the kinetic molecular theory is that gases diffuse into each other to form homogeneous mixture in which you cannot distinguish components; like in our air we cannot see nitrogen or oxygen gases separately.

4. c. *Students know* how to apply the gas laws to relations between the pressure, temperature, and volume of any amount of an ideal gas or any mixture of ideal gases. The Ideal Gas Law Equation: PV=nRT. A fixed number of moles *n* of gas can have different values for pressure *P,* volume *V,* and temperature *T in Kelvin.* Relationships among these properties are defined for an ideal gas and can be used to predict the effects of changing one or more of these properties and solving for unknown quantities. Students should know and be able to use the three gas law relationships Boyles Law $P_1V_1 = P_2V_2$., Charles Law $V_1/T_1 = V_2/T_2$, and Gay Lussac's Law $P_1/T_1 = P_2/T_2$.summarized in the Combined Gas Law Equation: $\frac{P1V1}{T1} = \frac{P2V2}{T2}$.

4. d. Students know the values and meanings of standard temperature and pressure (STP). Standard temperature is 273K (0°C) and standard pressure is 1 atmosphere (760 mm Hg). When volumes of gases are being compared, the temperature and pressure must be specified. For a fixed mass of gas at a specified temperature and pressure, the volume is also fixed.

4. e. Students know how to convert between the Celsius and Kelvin temperature scales. Some chemical calculations require an absolute temperature scale, called the Kelvin scale (K), for which the coldest possible temperature, absolute zero, is equal to 0 K. **There are no negative temperatures on the Kelvin scale.** In theory if a sample of any material is cooled as much as possible, the lowest temperature that can be reached is 0 K, experimentally determined as equivalent to −273.15°C. The Kelvin scale starts with absolute zero (0 K) because this is the theoretical lowest temperature limit. A Kelvin temperature is specified without the degree symbol. The magnitude of one unit of change in the K scale is equal to the magnitude of one unit of change on the °C scale.

4. f. Students know there is no temperature lower than 0 Kelvin. The kinetic molecular theory is the basis for understanding heat and temperature. The greater the atomic and molecular motion kinetic energy, the greater the observed temperature of a substance. If all atomic and molecular motion stopped, the temperature of the material would reach an absolute minimum. This minimum is absolute zero -273°C, or experimentally −273.15°C. The third law of thermodynamics states that this temperature can never be reached.

4. g.* Students know the kinetic theory of gases relates the absolute temperature of a gas to the average kinetic energy of its molecules or atoms. The value of the average kinetic energy for an ideal gas is directly proportional to its Kelvin temperature. Average kinetic energy can be related to changes in pressure and volume as a function of temperature. At 0 K all motion in an ideal monatomic gas ceases, meaning that the average kinetic energy equals zero.

4. h.* Students know how to solve problems by using the ideal gas law in the form PV = nRT. The relationships among pressure, volume, and temperature for a fixed mass of gas can be expressed as the ideal gas law, PV = nRT, where n represents moles and R represents the universal gas constant, which is 0.0821 liter-atmosphere per mole-Kelvin.

Interactive Notebook Score Sheet

Gas Laws Unit Spring Semester Quarter I

Quizzes/Formatives	**Date**	**Score/Max Score**	**Retake Needed (yes or no)**	**Peer Initial**	**Parent Initial**
Formative 21 Pressure, Temp, Vol Conversions					
Formative 22 Kinetic Molecular Theory					
Formative 23 Ideal Gas Law Quiz					
Formative 24 Combined Gas Law					
Unit 4 Test	1/29/14				

Name of Scored Assignment	Date Due	Score/Max	Peer Initials	Level of Effort

Histogram – Date (x axis) and progress on standards mastery with 5 advanced, 4 proficient, 3 basic, 2 below basis, and 1 incomplete (y axis)

5																
4																
3																
2																
1																
	1/6/2014	1/7/2014	1/8/2014	1/9/2014	1/10/2014	1/13/2014	1/14/2014	1/15/2014	1/16/2014	1/21/2014	1/22/2014	1/23/2014	1/24/2014	1/27/2014	1/28/2014	1/29/2014

Unit 5 Gases and their Properties Study Guide GUHS

California Chemistry Standard Set 4 - The kinetic molecular theory describes the motion of atoms and molecules and explains the properties of gases. *Textbook Chapters 3.1, 13 and 14*

Formative Assessments

1. Written/oral kinetic theory explanation of gas laws
2. Pressure and temperature conversions
3. Gas law problems (combined and ideal gas law)

Key Vocabulary Terms

1. kinetic molecular theory
2. kinetic energy
3. temperature
4. pressure
5. partial pressure
6. diffusion
7. random
8. Dalton's Law
9. Boyle's Law
10. Charles' Law
11. Gay-Lussac's Law
12. combined gas law
13. Avogadro's principle
14. molar volume
15. ideal gas law
16. ideal gas constant
17. STP
18. Absolute Zero
19. Intermolecular Forces
20. Van der Waals Forces
    a. Dispersion
    b. Dipole-Dipole
21. Hydrogen bonding

Concepts

1. The Kinetic Molecular Theory of Gases describes the behavior of molecules in a gas.
    a. Gases are small particles that are separated from one another by empty space. The volume of the particles is small compared with the volume of the empty space the particles occupy.
    b. There are no attractive or repulsive forces between the gas particles since they are so far apart.
    c. A gas consists of a collection of small particles traveling in constant random straight-line motion until a particle collides with another gas particle or with the walls of containers.
    d. Collisions between molecules are perfectly elastic. No energy is gained or lost during the collision; however kinetic energy can be transferred.
    e. Temperature is a measure of the average kinetic energy of the particles in a sample. At a given temperature all gases have the same kinetic energy. At absolute zero (0K) the kinetic energy is zero and all motion stops. Kinetic Energy (KE) = mv^2
2. Describe gases at the molecular level, the behavior of gases, and the measureable properties of gases.
3. Explain how motions and collisions of particles produce measureable properties such as pressure.
4. Compare and contrast states of mater: solid, liquid, gas in terms of kinetic energy and intermolecular forces.
5. Distinguish between homogeneous mixtures, heterogeneous mixtures, and pure substances.
6. Explain how temperature measure how hot a system is and is a measure of kinetic energy.

Items for Memorization –

- ✓ Recognize direct and inverse proportional relationships.
- ✓ Standard Temperature and Pressure (STP) occurs at 273K and 1 atmosphere
- ✓ Real gases do not behave like ideal gases when the pressure is extremely high (lots of them in a small space) and the temperature is extremely low (moving slow enough to notice each other)

Skills

1. Convert pressure units—kPa, mmHg, atm, psi, and others
2. Convert temperatures—ºC, K
3. Convert volumes—mL cm^3, L,
4. Solve algebraic gas law equations with several given quantities and one unknown variable.

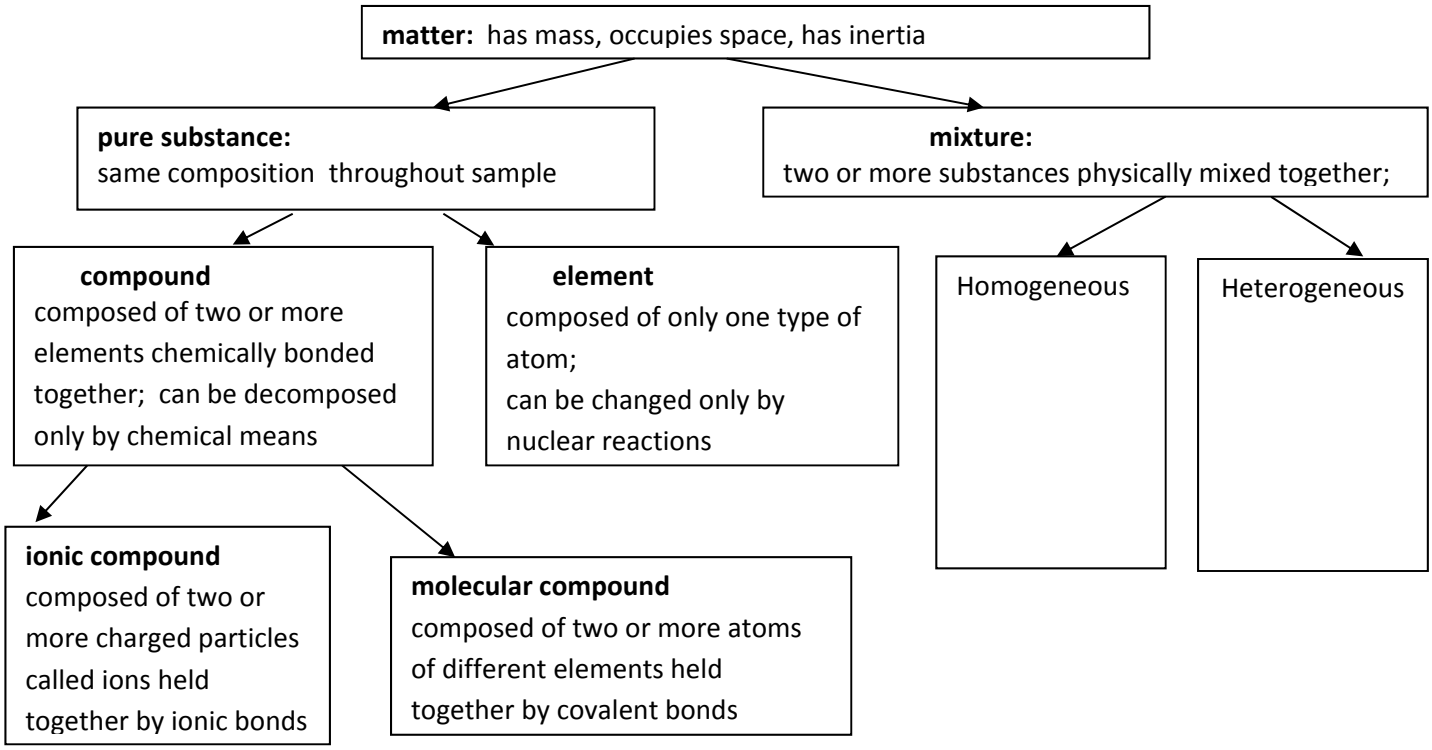
$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Gas Law	Fixed Values	Variable Relationships	Form for calculations
Boyles	n, T	Inverse	$P_1V_1=P_2V_2$
Charles	n, P	Direct	$V_1/T_1 = V_2/T_2$
Gay Lussac	n, V	Direct	$P_1/T_1 = P_2/T_2$

Where: n = number of moles, T = temperature (K), V = volume, P = pressure

Matter and Change

Chapters 3 and 13 in Textbook



Mixtures

A mixture is a combination of two or more pure substances in which each substance retains its individual properties. Concrete, most rocks, most metal objects, all food, and the air you breathe are mixtures that are often composed of many different substances. The composition of a mixture is variable. For example, the composition of salt water can be varied by changing the amount of salt or water in the mixture.

Two types of mixtures exist.

A heterogeneous mixture is one that is not blended smoothly throughout.

Examples of heterogeneous mixtures include smoky air in which suspended particles from the burn and gases from the burn and atmosphere are mixed or muddy water. You may have to use a magnifying glass or even a microscope, but if you can identify bits of one or more of the components of a mixture, the mixture is heterogeneous.

A homogeneous mixture is one that has a constant composition throughout.

By dissolving sugar in water, you create a homogeneous mixture. A homogeneous mixture is also called a solution. In solutions, the atoms and/or molecules of two or more substances are completely mingled with one another. Solutions do not have to be solids dissolved in liquids; they can be mixtures of various states of matter. For example, air is a gaseous solution containing nitrogen, oxygen, argon, carbon dioxide, water vapor, and small amounts of other gases. An alloy is a homogeneous mixture (solution) of two or more metals or of metals and nonmetals. Alloys are considered to be solid solutions.

1. Identify each of the following as an example of a homogeneous mixture or a heterogeneous mixture.

a. 70% isopropyl rubbing alcohol_____ b. a pile of rusty iron filings_____

c. concrete(cement rocks) _____ d. saltwater _____

e. gasoline _____ f. wheat bread _____

Physical means of separating a mixture include: filtration, evaporation, using known freezing points and boiling points to separate different liquids, distillation (boiling off the liquid to leave the solid component, and then condensing the vapor back to the liquid state).

1. The molecules are in constant motion. This motion is different for the 3 states of matter.

   - *Solid* - Molecules are held close to each other by their attractions of charge. They will bend and/or vibrate, but will stay in close proximity.

   - *Liquid* - Molecules will flow or glide over one another, but stay toward the bottom of the container. Motion is a bit more random than that of a solid.

   - *Gas* - Molecules are in continual straight line motion. A gas is composed of particles in constant motion. The kinetic energy of the molecule is greater than the attractive force between them (intermolecular forces), thus they are much farther apart and move freely of each other. Compared to the space through which they travel, the particles that make up the gas are so small that their volume can be ignored.

Physical states of matter:

- ✓ solid: particles packed very tightly together, particles are "fixed" in position relative to each other; lowest energy
- ✓ liquid: particles still very close together but particles can move around each other
- ✓ gas: particles very far apart from each other; highest energy

Under ordinary conditions, matter exists in three different physical forms called the states of matter—solid, liquid, and gas. Solid matter has a definite shape and a definite volume. A solid is rigid and incompressible, so it keeps a certain shape and cannot be squeezed into a smaller volume. A solid has these properties because the particles that make up the solid are packed closely together and are held in a specific arrangement. Liquid matter has a definite volume, like a solid, but flows and takes the shape of its container. A liquid is incompressible because its particles are packed closely. A liquid flows because the particles are held in no specific arrangement but are free to move past one another. Like a liquid, a gas flows and takes the shape of its container, but has no definite volume and occupies the entire space of its container. Gaseous matter has particles that are completely free to move apart to fill the volume of the container. Also, because of the space between its particles, a gas can be compressed to a smaller volume. A vapor is the gaseous state of a substance that is a liquid or a solid at room temperature.

Practice Problems. Identify as a solid, liquid, or gas.

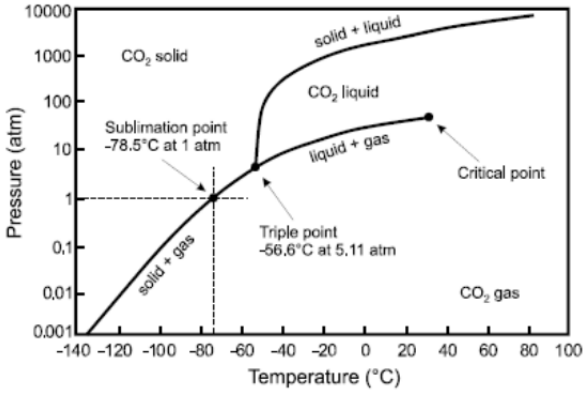
a. has a definite volume but flows_____

b. compressible _____

c. specifically arranged together_____

d. has a definite volume_____

e. always occupies the entire space of its container____



Pressure-Temperature phase diagram for CO_2.

Phase Changes

Most substances can exist in three states of matter—solid, liquid, and gas— depending on the temperature and pressure. States of substances are called phases when they coexist as physically distinct parts of a mixture, such as ice water. When Energy changes so do the phases.

Phase changes that require energy When you add ice to water, heat flows from the water to the ice and disrupts the hydrogen bonds that hold the water molecules in the ice together. The ice melts and becomes liquid. The amount of energy required to melt one mole of a solid depends on the strength of the forces keeping the particles together. The melting point of a crystalline solid is the temperature at which the forces holding the crystal lattice together are broken and the solid becomes a liquid. When liquid water is heated, some molecules escape from the liquid and enter the gas phase. If a substance is usually a liquid at room temperature (as water is), the gas phase is called a vapor. Vaporization is the process by which a liquid changes into a gas or vapor. When vaporization occurs only at the surface of a liquid, the process is called evaporation. As temperature increases, water molecules gain kinetic energy. When the vapor pressure of a liquid equals atmospheric pressure, the liquid has reached its boiling point, which is 100°C for water at sea level. The process by which solids change directly into a gas without first becoming a liquid is called sublimation, examples include: solid air fresheners and dry ice (CO_2).

Practice Problems. Classify each of the following phase changes.

a. dry ice to carbon dioxide gas _____

b. ice to liquid water _____

c. liquid bromine to bromine vapor _____

d. moth balls giving off a pungent odor_____

Phase changes that release energy Some phase changes release energy into their surroundings. For example, when a vapor loses energy, it may change into a liquid. Condensation is the process by which a gas or vapor becomes a liquid. It is the reverse of vaporization. Water vapor undergoes condensation when its molecules lose energy, their velocity slows, and hydrogen bonds begin to form between them. When hydrogen bonds form, energy is released. When water is placed in a freezer, **heat is removed** from the water. The water molecules **lose kinetic energy**, and their velocity decreases. When enough energy has been removed, the hydrogen bonds keep the molecules frozen in set positions. The freezing point is the temperature at which a liquid becomes a crystalline solid. When a substance changes from a gas or vapor directly into a solid without first becoming a liquid, the process is called deposition. Deposition is the reverse of sublimation. Frost is an example of water deposition.

Practice Problems Classify each of the following phase changes.

a. liquid water to ice _____

b. water vapor to ice crystals_____

c. water vapor to liquid water_____

d. water beads on the outside a cold drink glass_____

Intermolecular Forces

Van Der Waals	
London Dispersion Force	Nonpolar molecules. Ex. CH_4 N_2
Dipole-Dipole	Polar Molecules. Ex H_2S, SO_2
Hydrogen Bonding	H-F, H-O-, H-N-, NH_3, H_2O, amines NH_2, and alcohols C-OH
Metallic	Metals, Ag, Pb
Ionic – crystalline structures	Salts, NaCl, $CaCO_3$ (note: "ates" contain covalent bonds)
Covalent network	C (graphite), C (diamond), SiO_2, (Note: graphite = London too)

13.2 Forces of Attraction

The attractive forces that hold particles together in ionic, covalent, and metallic bonds are called intramolecular forces. Intermolecular forces, which are weaker than intramolecular forces, also can hold particles, bonded or not, together. Three types of intermolecular forces are described below: dispersion forces, dipole–dipole forces, and hydrogen bonds.

Dispersion forces Weak forces that result from temporary shifts in the density of electrons in electron clouds are called **dispersion forces,** or London forces. When two nonpolar molecules are in close contact, the electron cloud of one molecule repels the electron cloud of the other molecule. As a result, the electron density in each electron cloud is greater in one region of the cloud. Two temporary dipoles form. Weak dispersion forces exist between oppositely charged regions of the dipoles. Dispersion forces, which are the weakest intermolecular forces, are important only when no stronger forces are acting on the particles. Dispersion forces are noticeable between identical nonpolar molecules as the number of electrons involved increases. For example, an increase in dispersion forces explains why fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid at room temperature. As the number of electrons increases, the dispersion force increases causing molecules to get closer and closer, moving from the gaseous state to solid state.

Dipole–dipole forces Attractions between oppositely charged regions of polar molecules are called **dipole–dipole forces.** Polar molecules have a permanent dipole and orient themselves so that oppositely charged regions match up. Dipole–dipole forces are stronger than dispersion forces as long as the molecules being compared are similar in mass.

Hydrogen bonds A **hydrogen bond** is a dipole–dipole attraction that occurs between molecules containing a hydrogen atom bonded to a small, highly electronegative atom with at least one lone electron pair. The hydrogen must be bonded to a fluorine, an oxygen, or a nitrogen atom. Hydrogen bonds explain why water is a liquid at room temperature, while compounds of comparable mass are gases.

Read Page 393-395 in textbook and answer the following questions:

1. Compare Intermolecular Forces with Intramolecular Forces

_____

_____

_____

2. What are dispersion Forces and why are they weaker than Dipole-Dipole Forces

_____

_____

_____

3. Hydrogen bonds are a special type of dipole-dipole attraction that occurs between hydrogens bonded to electronegative atoms. These hydrogen bonds are important for the conformation (shape) of many biological macromolecules like DNA and proteins. Explain the difference between hydrogen bonds in water and ammonia.

_____

_____

_____

Kinetic Theory of Gases

The theory is a model or a mental picture that enables us to better understand our observations. The theory attempts to elucidate the behavior of gases is known as kinetic theory of gases.

Assumptions or postulates of the kinetic molecular theory of gases are given below:

(i) Gases consist of large number of identical particles that are so small and so far apart on the average that the actual volume of the molecules is negligible in comparison to the empty space between them. They are considered as point masses. This assumption explains the great compressibility of gases.

(ii) There is no force of attraction or repulsion between the particles of a gas at ordinary temperature and pressure because the gas particles are far apart. The support for this assumption comes from the fact that gases expand and occupy all the space available to them.

(iii) Particles of gas are always in constant and random motion. If the particles were in rest and occupied fixed position, then a gas would have had a fixed shape which is not observed. Particles of gas move in all possible direction in straight lines. During their random motion, they collide with each other and with the walls of container. Pressure in exerted by the gas as a result of collision of the particles with the walls of the container

(iv). Collision of gas molecules are perfectly elastic. This means that total energy of molecules before and after the collision remains same although energy can be transferred..

(v) Temperature is a measure of the average kinetic energy of the particle in a sample. At a given temperature all gases will have the same average kinetic energy regardless of mass. K.E. = ½ mv2 so heavier particles will move slower even though they have the same kinetic energy. At absolute zero (0K) the kinetic energy is zero and all motion stops.

13.1 Gases

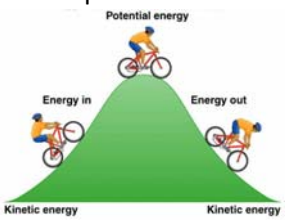
In the late 1800s, two scientists, Ludwig Boltzmann and James Maxwell, independently proposed a model to explain the properties of gases in terms of particles in motion. This model is now known as the **kinetic-molecular theory.**

The model makes the following assumptions about the size, motion, and energy of gas particles.

- **Particle size** The particles in a gas are separated from one another by empty space. The volume of the empty space is much greater than the volume of the gas particles themselves. Because gas particles are so far apart, there are no significant attractive or repulsive forces between them.
- **Particle motion** Gas particles are in constant, random motion. Until they bump into something (another particle or the side of a container), particles move in a straight line. When gas particles do collide with something, the collision is said to be elastic. An **elastic collision** is one in which no kinetic energy is lost. Although kinetic energy may be transferred from one particle to another, the total amount of kinetic energy of the two particles does not change.
- **Particle energy** Mass and velocity determine the kinetic energy of a particle, as represented in the equation below.

$$KE = \frac{1}{2} mv^2$$

KE = kinetic energy
m = mass of the particle
v = velocity of the particle



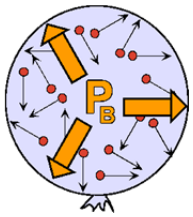
The velocity of a particle includes both its speed and its direction. Each particle in a sample containing only one gas will have the same mass but not the same velocity. Thus, all the particles in a sample of gas do not have the same kinetic energy.

Temperature is a measure of the *average* kinetic energy of the particles in a sample of matter.
At a given temperature, all gases have the same average kinetic energy.

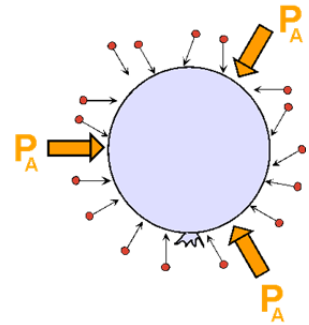
EXPLAINING THE BEHAVIOR OF GASES The kinetic-molecular theory explains the following behavior of gases.

- **Low density.** Density is a measure of mass per unit volume. The difference between the high density of a solid and the low density of a gas is due mainly to the large amount of space between the particles in the gas. There are fewer particles in a gas than in a solid of the same volume.
- **Compression and expansion.** A gas will expand to fill its container. Thus, the density of a sample of gas will change with the volume of the container it is placed in. The gas will become more dense as it is **compressed** into a smaller container. The gas will become less dense as it **expands** in a larger container.
- **Diffusion .** Gas particles flow past each other easily because there are no significant forces of attraction between them. **Diffusion** refers to the random movement of one material through another, such as when one gas flows into a space already occupied by another gas. The rate of diffusion depends mostly on the mass of the particles. **Lighter particles diffuse more quickly than heavier particles.** Because lighter particles have the same average kinetic energy as do heavier particles at the same temperature, lighter particles must have, on average, a greater velocity.
- **Effusion.** If you have ever seen a tire deflate from a puncture, you are familiar with effusion. Effusion is the escape of a gas through a small opening in its container.

Gas Pressure



Gas molecules inside a volume (a balloon) are freely and constantly moving around. During this molecular motion they frequently collide with each other and with the surface of any enclosure there may be. In a small balloon that would be many thousands of billions of collisions each second. The internal gas pressure in a balloon, P_B, is caused by the impacts of moving gas molecules, as they collide with the skin of the balloon from the inside.



The force of impact of a single one collision is too small to be sensed or measured. However, taken all together, this large number of impacts of gas molecules exerts a considerable force onto the surface of the enclosure: the gas pressure. **Thegreater the number of collisions per area of enclosure, the greater the pressure.** Collisions create pressure

Explain why if I let air out of the balloon the pressure inside decreases

_____

_____

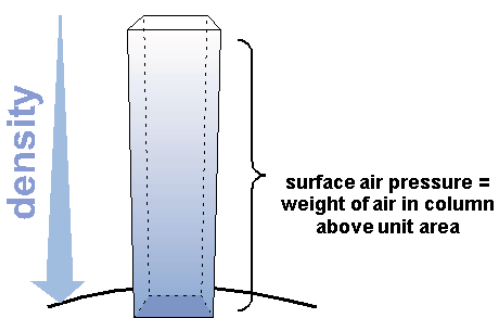
_____

Atmospheric Pressure – Pressure Profile

In the example of the balloon (above), there is not only gas inside the balloon (exerting pressure from the inside), but there is also gas (air) on the outside, exerting pressure onto the outside surface of the balloon The atmospheric pressure outside a balloon, P_A, is given by the impacts of moving gas molecules, as they collide with the skin of the balloon from the outside. The rate, at which the skin of the balloon is bombarded by air molecules, is dependent on how tightly the gas molecules are packed, or on the **gas density**:

$$(\text{Gas})\ \text{Density} = \frac{\text{Mass (of gas)}}{\text{Volume}}$$

Since gas is compressible, its density depends on the force that is used to compress it. In the atmosphere, the force that compresses the air at the surface is just the weight of all the air in the atmospheric column above it.



Surface pressure and density in an air column

<u>At the surface the atmospheric pressure is on average P_0 = 101.3 kPa</u>

The higher we go in the atmosphere, the less air remains in the column above us. Thus, the atmospheric pressure always decreases with height. Similarly, air density decreases with height, because the overload to compress the air gets less and less, as we go higher.

$$\text{Atmospheric Pressure} = \frac{\text{Weight of Overlying Air}}{\text{Unit Area}}$$

At the surface, the air is densest. Thus, as we rise from the surface up through the first kilometer of the atmosphere, we leave a lot of dense air below us: the overload (and thus the density and pressure) decreases quickly. The reduction of pressure decreases for every consecutive vertical stretch of atmosphere because the overload gets less and less. By 6 km high the atmospheric pressure has about half of the mass of air overlying. Thus, the atmospheric pressure at 6 km height is only half that at the surface (i.e., P_{6km} = 50.5 kPa). Between 6 and 12 km the overload (and the pressure) is halved again: only ¼ of the surface pressure is left, and this halving occurs every 6 km.

Diagram air pressure changes moving up a mountain. **Sketch a graph of elevation vs pressure**

Gas pressure Units of Measure When gas particles collide with the walls of their container, they exert pressure on the walls. **Pressure** is force per unit area.

The pressure exerted by the particles in the atmosphere that surrounds Earth is called atmospheric pressure, or air pressure. Air pressure varies at different locations on Earth. At Earth's surface, air pressure is 101.3 kPa. Air pressure at higher altitudes, such as on a mountaintop, is slightly lower than air pressure at sea level.

How Much is a Pascal (Pa)
The pascal (Pa) is the SI unit of pressure. One **pascal** is equal to a force of one newton per square meter. Some scientists use other units of pressure. For example, engineers use pounds per square inch (psi). Barometers and manometers measure pressure in millimeters of mercury (mm Hg). A unit called the torr is equivalent to mm Hg. Air pressure is often reported in a unit called an atmosphere (atm). One **atmosphere** is equal to 760 mm Hg, 760 torr, or 101.3 kilopascals

Pressure Conversions

Pressure is defined as the force pushing over a certain area. A gas pressure results from the many collisions between gas particles and their container. One common unit for pressure is the newton per square meter (N/m^2). This unit is called a **pascal** (Pa). Since it is so small it is often reported in thousands of pascals or **kilopascals** (kPa). The gases surrounding the earth exert a pressure of approximately 1 **atmosphere** (atm) at sea level. There are other units used to measure pressure (see the table below and refer to CST Reference Sheet).

Example:

How many atm are in 1520 mmHg?

$$P_{atm} = \frac{\cancel{1520}\ 2\ \cancel{mmHg}}{} \left| \frac{1\ atm}{\cancel{760\ mmHg}\ 1} \right. = \boxed{2\ atm}$$

How many atmospheres (atm) are in 7.35 pounds per square inch ($lbs./in^2$)?

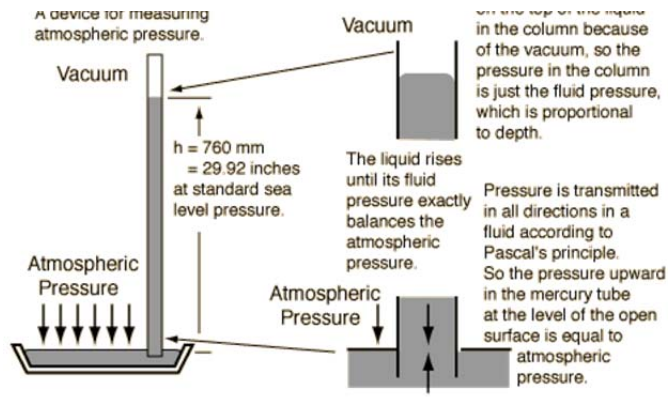
How many atmospheres (atm) are in 202.6 kilopascals (kPa)?

AIR PRESSURE Measurements

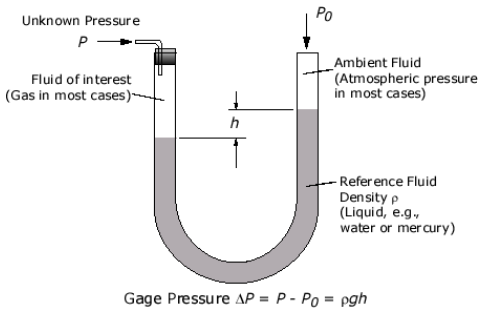
Air pressure is measured using a **barometer.** A barometer consists of a thin tube closed on one end and filled with mercury. The tube is placed so that the level of the mercury is determined by air pressure. The mercury rises when the air pressure increases and falls when the air pressure decreases.

A **manometer** is an instrument used to measure gas pressure in a closed container. A flask containing gas is attached to a sealed U-shaped tube that contains mercury. The mercury is level across the two arms of the U. When a valve between the flask and the tube is opened, gas particles diffuse into the tube and push down on the mercury. The difference in the height of the mercury in the two arms of the U is used to calculate the pressure of the gas in the flask.

Barometer



Manometer



Hydrostatic Pressure – Mercury Barometer

The principle that the pressure at a given level is equivalent to the weight of the overlying column is not only true for air, but for fluids (gases and liquids) in general. The pressure generated by an overlying column of fluid is thus termed the hydrostatic pressure.

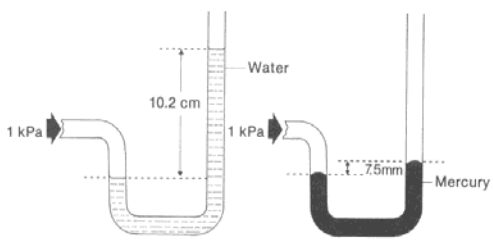


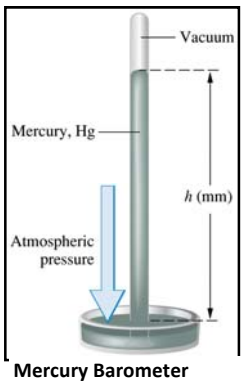
Figure 1.11 Comparison of water and mercury manometers.

If a much heavier liquid substance is used to balance this air column, only a relatively small length would be needed. In addition, because the density of liquids does not change with height (most liquids are incompressible), such an equivalent liquid column has a well defined upper boundary (below a vacuum). One of the heaviest liquids at room temperature is mercury (Hg) and the height of the Hg-column that is equivalent to normal pressure (101.3 kPa) is only 760 mm Hg long (29.92 inches Hg). For this reason, columns of mercury, "hanging" in an inverted vacuum tube, can be used as practical instruments to measure atmospheric pressure.

If water were used instead of mercury, the height of the column equivalent to normal pressure would be 10.33 m - not a very practical length of tube to work with, due to water's lower density.

Show what a barometer would look like:

Lower Pressure, higher altitude



Higher Pressure, below sea level

Mercury Barometer
At sea level, h = 760 mm Hg

SCUBA SCIENCE

1. What is meant by scuba?

2. Who invented scuba?

3. How did the invention of scuba increase the maneuverability and convenience of diving?

4. Why does the pressure acting on your body increase when you descend in the sea?

5. Why is it necessary to use a scuba when diving?

6. Explain how the lungs react to an increased pressure when a person dives.

7. Explain why is an understanding of partial pressure is important in scuba diving.

8. Explain how bends occur. How can it be prevented from happening?

9. What are the symptoms of nitrogen narcosis?

10. How could nitrogen narcosis be prevented?

In 1943 Jacques-Ives Cousteau and Emile Gagnan invented scuba: a Self-Contained Underwater Breathing Apparatus. No longer did divers need to be tethered to the surface by air hoses. The increased maneuverability and convenience gave rise to the sport of scuba diving. It also greatly increased the use of diving for scientific research. All divers must have a good understanding of the science involved in diving in order to dive safely.

At the water's surface, the pressure on your body due to the mass of air around you is 1 atmosphere. Under water, the pressure increases due to the added mass of water. Every 10 meters of depth adds I atmosphere pressure. Thus the total pressure on your body at a depth of 10 meters will be 2 atm, at 20 m 3 atm, and so on. At around 40 m the pressure on your chest would make it impossible for you to inflate your lungs to breathe. If the pressure in your lungs had also increased as you went down, however, you would be able to breathe normally. Scuba equipment provides air to lungs at a pressure to match that of the underwater environment. This enables the diver to breathe comfortably.

A diver at 20 m is under a pressure of 3 atmospheres. The scuba equipment is maintaining the same pressure in his or her lungs. The average lung capacity of a human being is 6-7 liters. According to Boyle's law, this amount of air will expand three times its volume if the pressure is reduced to 1 atmosphere. If a diver ascends to the surface without exhaling steadily along the way, the air held in the lungs will expand as the pressure drops. The increase in volume can rupture the lungs.

An understanding of partial pressure is also important in scuba diving. When the pressure on a mixture of gases increases, the partial pressure of each of the gases increases proportionately. This means that in the compressed air that a diver breathes, every ingredient is at higher pressure. At depths of 30 meters the partial pressure of carbon dioxide in air is sufficient to poison a diver. Even oxygen can be toxic if the total gas pressure is 2 atmospheres. At depths over 10 meters the length of the dive becomes very important in preventing such toxic effects.

Decompression sickness, or "bends," is explained by another important gas law. The solubility of a gas in liquid is proportional to the pressure of the gas above the liquid. A diver is breathing air at higher than atmospheric pressure. Thus the nitrogen in the air will be more soluble in his or her blood. When the pressure drops as the diver ascends, the nitrogen again become less soluble. If the drop in the pressure occurs too rapidly, the nitrogen will come out of the blood to form tiny bubbles. This effect is just like the bubbles that form in carbonated drink when you open the cap and relieve the pressure in the bottle. If the nitrogen bubbles form in the joints or muscles, they cause a great deal of pain. If they in the spinal cord, brain, or lungs they can cause paralysis or death. Ascending from the dive slowly can prevent decompression sickness. If that is not possible, the diver can be brought to the surface rapidly and put in the recompression chamber. Here he or she is recompressed to a pressure of 6 atmospheres. Then the pressure is gradually reduced so that nitrogen can be eliminated through lungs.

Air cannot be used in underwater breathing apparatus at depths greater than about 45 meters because of nitrogen's narcotic effects. Nitrogen narcosis or "rapture of the deep" occurs at depths greater than 30 meters. The symptoms are similar to intoxication by alcohol: feelings of happiness and overconfidence, tingling or numbness in the arms or legs, and memory impairment.

To prevent narcosis, divers exploring the ocean floor or workers building tunnels breathe a mixture of gases that does not contain nitrogen. A diving mixture at 300 meters may contain 72% neon, 24% helium, and 4% oxygen, for example. Moreover, because neon and helium are less soluble in blood than nitrogen, such a mixture allows more rapid decompression and reduces danger of bends.

Conversions to Solve Gas Law Problems

When solving Gas Law Problems it is very important that the units match. All temperatures used for Gas Law Problems are in Kelvin.

1. Convert these Temperatures to Kelvin (K) or Celsius (°C). Show ALL Work

25°C

-27°C

100 K

273 K

2. Make the following Pressure Conversion

760 mm Hg to atm

800 mm Hg to KPa

380 mm Hg to psi

0.75 atm to mm Hg

0.25 atm to KPa

3. Convert the following volume measures

150 mL to L

6400 cc to L

43 L to mL

125 L to mL

3.25 L to mL

4.5 mL to L

Chemistry Benchmark: Unit Conversions for the Gas Laws Standard Check

Directions: Complete the following tables, showing your work for each lettered box beside the corresponding letter below. Include units on your work, and write your final answers in the tables.

TEMPERATURE	
K	°C
373 K	(A)
(B)	56°C
(C)	154°C
128 K	(D)

PRESSURE		
mm Hg	kPa	atm
760 mm Hg	(E)	(F)
(G)	151.8 kPa	(H)
(I)	(J)	0.5 atm
3040 mm Hg	(K)	(L)

(A)

(B)

(C)

(D)

(E)

(F)

(G)

(H)

(I)

(J)

(K)

(L)

Volume Benchmark Convert the following volume measures

300 mL to L

0.25 L to mL

6.2 L to cc

1.3 L to cc

Ideal Gas Law and Real Gases

The Ideal Gas Law [1]

Key Concepts

An Ideal Gas obeys the Ideal Gas Law exactly. **PV = nRT**

where

P=pressure
V=volume
n=moles of gas
T=temperature Temperature is in Kelvin(K)
R = gas constant (depends on the units of pressure and volume)

R = 8.314 L kPa K^{-1} mol^{-1} (8.314 L·kPa/K·mol) if
Pressure is in kilopascals(kPa) Volume is in liters (L) Temperature is in Kelvin (K)

R = 0.0821 L atm K^{-1} mol^{-1} (0.0821 L·atm/K·mol) if
Pressure is in atmospheres(atm) Volume is in liters(L) Temperature is in Kelvin(K)

An Ideal Gas is modeled on the Kinetic Theory of Gases which has 5 basic postulates:

a. Gases consist of small particles (molecules) which are in continuous random motion
b. Intermolecular forces are negligible – No repulsion or attraction
c.Pressure is due to the gas molecules colliding with the walls of the container
d. Collisions are elastic. The total energy is unchanged but energy can be transferred.
e. Temperature is a measure of average kinetic energy

Real Gases deviate from Ideal Gas Behavior because:

- ✓ at low temperatures the gas molecules have less kinetic energy (move around less) so they *do* attract each other
- ✓ at high pressures the gas molecules are forced closer together so that the volume of the gas molecules becomes significant compared to the volume the gas occupies

Under ordinary conditions, deviations from Ideal Gas behavior are so slight that they can be neglected. A gas which deviates from Ideal Gas behavior is called a *non-ideal gas*.

Ideal Gas Law Calculations

Calculating Volume of Ideal Gas: What volume is needed to store 0.050 moles of helium gas at 202.6kPa and 400K?

PV = nRT

Summary of Variables

P = 2 atm
n = 0.050 mol
T = 400K
V = ? L
R = 0.0821 L atm K^{-1} mol^{-1}

Rearranging the equation to solve for V

[1] http://www.ausetute.com.au/idealgas.html

Calculating Pressure of Ideal Gas: What pressure in atm will be exerted by 20.16g hydrogen gas in a 15L cylinder at 27°C? HINT to find moles of gas convert g H_2 to moles H_2. Select which universal gas constant will you use?

PV = nRT

Summary of Variables

Rearranging the equation to solve for P

Calculating moles of gas: A 50L cylinder is filled with argon gas to a pressure of 10130.0 kPa at 30ºC. How many moles of argon gas are in the cylinder?

PV = nRT

Summary of Variables

Rearranging the equation to solve for n

What is the mass of argon contained in the 50-L cylinder

Calculating gas temperature: To what temperature does a 250mL cylinder containing 0.40g helium gas need to be cooled in order for the pressure to be 101kPa?

PV = nRT

Summary of Variables

Rearranging the equation to solve for T

What differentiates a real gas from an ideal gas? The Ideal Gas Law assumes[2]:

- ✓ The molecules of an ideal gas have no volume
- ✓ And, there are no attractive forces between the molecules within a gas.
- ✓ The gas molecules move in a random manner, and their collisions with each other, and the container, are perfectly elastic.

This, of course, is not true. There are attractive forces between the molecules, and the gas molecules do have volume, although small. With these in mind, Real gases behave like Ideal Gases at the following conditions:

- At low pressures: the gas molecules spread apart so that the small volume of each gas molecule is negligible relative to the entire volume of the gas.
- At high temperatures: the gas molecules are moving so quickly; the small attractive forces between the molecules of a gas are overcome by their rapid speed.

Ideal Gas or a Real Gas

Identify the following as characteristic of an Ideal Gas (I) or a Real Gas (R)

_____ a. Particles are very small

_____ b. Particles have no volume

_____ c. Will have 0 volume at 0 K

_____ d. Would be a solid at 0 K

_____ e. Have weak attractive forces

_____ f. Have no attractive forces

_____ g. Obey gas laws at high T, low P

_____ h. Obey gas laws at all T and P

_____ i. Would never become a liquid

_____ j. Will eventually become a liquid

Which real gas would act "most" like an Ideal gas? Why?

If a real gas has *strong* attractive forces, would its Volume at *Low Temperature* be *more* or *less* than an ideal gas? Why do you think so?

[2] http://aspire.cosmic-ray.org/javalabs/java12/gaslaws/index.htm

Ideal Gas Law Problems (show all work

1. What is the pressure of 30.0 moles of N_2 gas in a 500.0 L storage tank at 27.0°C?

2. What is the volume occupied by 8.00 grams of O_2 gas at STP? (change g-->mol)

3. What is the temperature of methane within a container that has 16.42 L of gas at 6.0 atm pressure and contains 1.5 moles of carbon dioxide?

4. Calculate the volume (in liters) occupied by 16.0g of O2 at a pressure of 150kPa and 300K?

5. A gas in a 500 mL storage tank has a pressure of 1520 mm Hg at 27.0°C.
   a. How many moles of a gas are in this tank?

   b. If this gas is methane, CH_4, what is the mass of the CH_4 gas in the tank?

Gas Law Permutations *IN PENCIL*

To "Solve for" something means that the variable must appear all by itself ABOVE a fraction bar on either side of an equal sign.

Solve the gas law equations for each of the variables listed in the boxes
The Ideal Gas Law where P is pressure, V is volume, n is number of moles, R is the gas constant and T is temperature (K)

$$PV=nRT$$

Solve for P	Solve for n

V	R

T	Calculate R when n = 1mole, V = 22.4 L, T = 273K, and P = 1 atm

Explain how to solve an Ideal Gas Law Problem

The Combined Gas Law where P is pressure, V is volume, n is number of moles is constant, and T is temperature (K)

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

P1	P2	T1
V1	V2	T2

Explain how to solve combined Gas Law Problem

The Gas Laws

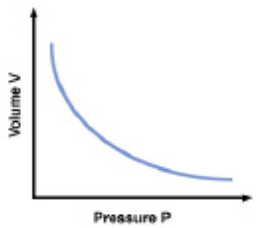
The example of the gas-filled balloon can also be used to explore the basic gas laws . In the following, assume that the balloon is tight, so that the amount or mass of air in it stays the same: $\mathbf{m_a = const}$. With density being the ratio of mass per volume, the gas density of the balloon thus varies only with its volume (when mass is held constant).

If we squeeze the balloon, we compress the air and two things will happen:

- the air pressure in the balloon will increase.
- the density of the air in the balloon will increase.

Since density is mass over volume, and the mass stays constant, the rise in density means that the volume of the balloon decreases: pressure goes up; volume goes down. This finding is expressed more precisely by Boyle's Law

Boyle's Law (Robert Boyle, Anglo-Irish scientist 25 January 1627 – 31 December 1691)



Boyle's law states that, *at a constant temperature, the volume of a given mass of gas varies inversely with pressure.* For two states of pressure (P_1, P_2) and two corresponding volumes (V_1, V_2), this is stated mathematically:

$$P_1V_1 = P_2V_2$$

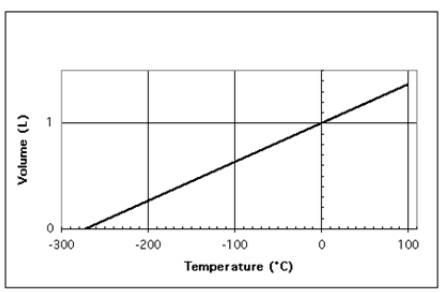
Predict what would happen to pressure if I increase volume

_____

_____

Charles's Law (**Jacques Charles**, French scientist November 12, 1746 – April 7, 1823)

Charles' law (also known as the law of volumes) is an experimental gas law which describes how gases tend to expand when heated. It was first published by French natural philosopher Joseph Louis Gay-Lussac in 1802, although he credited the discovery to unpublished work from the 1780s by Jacques Charles.



By warming the balloon up, we increase the speed of the moving gas molecules inside it. This in turn increases the rate at which the gas molecules bombard the skin of the balloon. Because the balloon's skin is elastic, it expands upon this increased pushing from inside, and the volume taken up by the same mass of gas increases with temperature. In consequence, the density [density =mass/volume] decreases with rising temperature. Cooling the balloon down again will make the balloon shrink. Thus Charles's law states that *at a constant pressure, the volume of a given mass of gas is directly proportional to its (absolute) temperature*. It must be noted that in this case (and *whenever temperature appears in a multiplication or a division*) the ***absolute or Kelvin scale must be used*** for temperature.

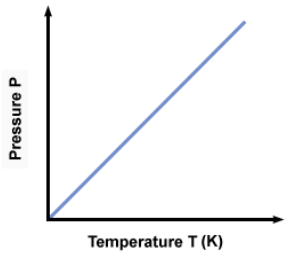
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Predict what would happen to volume if I incresed temperature

_____

_____

Louis Joseph Gay-Lussac[3] French Scientist (6 December 1778 – 9 May 1850) He received his early education at the hands of the Catholic Abbey of Bourdeix. Among Gay-Lussac's early work was an extensive investigation of how the volume of various gases changes with temperature. The English scientist John Dalton was independently studying the same phenomenon. Both found that the volume V of all gases studied increased similarly with higher temperature T when pressure P was held constant (VαT at constant P). In 1787. In his own scientific memoirs Gay-Lussac acknowledged hearing of Charles's work. Thus, the law governing the thermal expansion of gases based on Pressure instead of Volume is attributed to Gay-Lussac as a nod to his contribution.



Gay-Lussac's Law: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ note volume is held constant

Predict what would happen to pressure if I decreased temperature

_____

_____

Combined Gas Law

The Combined Gas Law is a combination of Boyle's, Charles' and Gay Lussac's Laws. The Combined Gas Law describes the relationship between pressure, volume, and temperature. For example, if the pressure increased, wither the volume would decrease or the temperature would increase. The Combined Gas Law can be used to solve any gas problem in which the number of moles remains constant. If one of the Combined Gas Law variables remains constant, cross it out of your equation or setting it to the same value on either side of the equals sign.

Here is one way to "derive" the Combined Gas Law[4]:

Step 1: Write Boyle's Law:

$P_1V_1 = P_2V_2$

Step 2: Multiply by Charles Law:

$P_1V_1^2 / T_1 = P_2V_2^2 / T_2$

Step 3: Multiply by Gay-Lussac's Law:

$P_1^2V_1^2 / T_1^2 = P_2^2V_2^2 / T_2^2$

Step 4: Take the square root to get the combined gas law:

$P_1V_1 / T_1 = P_2V_2 / T_2$

Sample Problem Video: http://www.wisc-online.com/objects/ViewObject.aspx?ID=GCH5404

[3] http://www.chemistryexplained.com/Fe-Ge/Gay-Lussac-Joseph-Louis.html

[4] http://www.chemteam.info/GasLaw/Gas-Combined.html

Combined Gas Law Practice I

1. A 3.00 L pocket of air at sea level has a pressure of 100. kPa. Suppose the pocket rises in the atmosphere to a certain height and expands to a volume of 10.0 L. What is the pressure of the air at the new volume?

2. A weather balloon has a volume of 1750 L at 105 kPa. The balloon is released into the atmosphere. At the highest point above the ground, the pressure on the balloon is 35.0 kPa. What is the new volume of the balloon at this new height?

3. A ball has a volume of 5.00 L at a pressure of 100. kPa at the surface of the ocean. What is the volume of the ball if it is submerged 100. m below the surface? The pressure increases by 9.80 kPa for every meter of ocean depth.

4. Determine the pressure change when a constant volume of gas at 1.00 atm is heated from 20.0 °C to 30.0 °C.

5. A gas has a pressure of 0.323 atm at 50.0 °C. What is the pressure at standard temp?

6. A gas has a pressure of 700.0 mm Hg at 77.0 °C. What is the temperature at standard pressure?

7. The gas in a balloon occupies 3.00 L at 300. K. At what temperature will the balloon expand to 9.00 L?

8. A 250. mL volume of gas is collected at 330. K. What volume would the sample occupy at 27°C?

9. A flexible container holds 45.0 L of air at 298 K. What would be the volume of the container if it were cooled to -124°C?

CHEMISTRY: COMBINED GAS LAW PRACTICE II *Solve the following problems. Show your work and units.*

1. A gas has an initial volume of 15 L and pressure 2 atm. If the temperature increases from 330 K to 462 K and the pressure reduced to 1 atm, find the new volume.

2. 2L of gas exerts 1.2 atm of pressure. If the temperature is raised from 27°C to 600 K and the volume increased to 4000mL, find the new pressure.

3. A sample of oxygen takes up 3400 cm^3 of space when it is under 505 kPa of pressure. When the pressure is changed to 3404 kPa, find the new volume at constant temperature.

4. The pressure and temperature of some N_2 gas drops from 303 kPa at 300 K to 202 kPa at -73°C. If the initial volume is 2 L, find the new volume.

5. The pressure of 500 mL neon changes from 760 mm Hg to 1520 mm Hg 1.5 L. If the initial temperature -73°C, what is the new temperature (in °C)?

6. When the temperature of a gas changes, its volume increases from 12 cm^3 to 36 cm^3. If the **final** temperature is measured to be 500 K, what was the initial temperature (in °C) if constant pressure was maintained?

7. The temperature of a sample of gas in a steel container (no volume change) at 30.0 kPa is increased from -127.0 °C to 1.50×10^3 K. What is the final pressure inside the tank?

8. Calculate the final pressure inside a scuba tank after it cools from 1.50×10^3 K to 27.0 °C. The initial pressure in the tank is 130.0 atm.

9. A gas at STP occupies 28 cm^3 of space. If the pressure changes to 4 atm and the temperature increases to 373°C, find the new volume.

Discover What Does a Graph of Pressure and Volume Show?

Volume (mL)	Pressure (kPa)
100	60
90	67
80	75
70	86
60	100

1. In an experiment, the volume was varied for a constant temperature of gas. Gas pressure was measured after each 10 mL change.
2. Show volume on the horizontal axis (x) with a scale from _____ to _____. Show pressure on the vertical axis (y) from _____ to _____.
3. For each pair measurements, **draw a point on the graph.**
4. D**raw a line of best fit** between the points.
5. Which Gas Law does this represent?_____

What is the relationship between volume and pressure of gas when the temperature is held constant? Use the graph to describe the relationship between the volume and pressure of a gas, when temperature is held constant.

_____

_____

_____

_____

Discover What Does a Graph of Pressure and Temperature Show?

1. In an experiment, the temperature was varied for a constant volume of gas. Gas pressure was measured after each 5° C change. **Convert to Kelvins.**
2. Show temperature on the horizontal (x) axis with a scale from _____ to _____. Show pressure on the vertical (y) axis from _____ to _____. Label axes.
3. For each pair measurements, **draw a point on the graph.**
4. D**raw a line of best fit** between the points.
5. Which Gas Law does this represent?_____

Temperature (°C)	Temperature (K)	Pressure (kPa)
0		8
5		11
10		14
15		17
20		20
25		23

What is the relationship between pressure and temperature of gas when the volume is held constant? Use the graph to describe the relationship between the pressure and temperature of a gas, when volume is held constant.

Discover What Does a Graph of Temperature and Volume Show?

1. In an experiment, the temperature was varied for a constant pressure of gas. Volume measured after each 10 K change.
2. Show temperature on the horizontal axis with a scale from _____ to _____. Show volume on the vertical axis from _____ to _____.
3. For each pair measurements, **draw a point on the graph.**
4. D**raw a line of best fit** between the points.
5. Which Gas Law does this represent?_____

Temperature (K)	Volume (mL)
273	50
283	52
293	54
303	56
313	58
323	60
333	62
343	63
353	66
363	67
373	69

What is the relationship between volume and temperature of gas when the pressure is held constant? Use the graph to describe the relationship between the volume and temperature of a gas, when pressure is held constant.

Define the following

Horizontal Axis

Vertical Axis

Manipulated Variable

Responding Variable

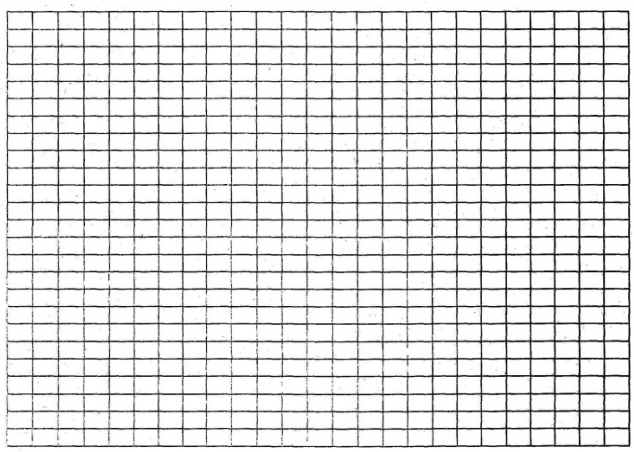
Linear Relationship

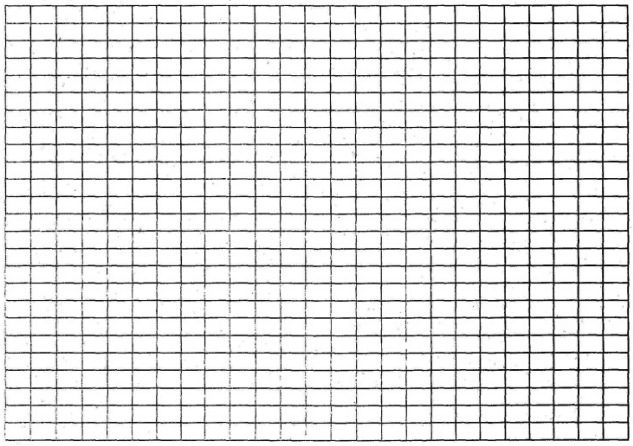
Non-linear Relationship

Directly Proportional

Varies Inversely

How can you tell the difference between a graph in which on variable is **directly proportional** to another and a graph in which two variables **vary inversely?**





Experimental Demonstration – Charles's Law

Objective: To determine the relationship between the volume of a confined gas and its temperature at the same pressure.

Materials:

Beakers (1000 mL and 600 mL)	balloon
Hotplate	thermometer

Hypothesis: _____

_____

_____

Procedure:

Half-fill the 1000 mL beaker with water. Inflate a balloon to about 2 inches diameter. Tie the balloon tightly to make sure that no air leaks.

Place the balloon into the water in the beaker. Put the 600 mL flask half filled with water over the balloon to completely submerge the balloon in the water.

Heat the system on the hotplate. Observe what happens to the balloon while monitoring the temperature of the water. Take note of any changes on the water level in the beaker and on the balloon.

Turn off the hotplate when the water temperature reaches 70^0C. Let the water cool. Observe the size of the balloon and the water level in the beaker.

Questions:

1. Describe what happens to the balloon and the water level while the beaker is heated.

2. Describe what happens to the balloon and the water level when the water cools.

3. What serves as the pressure in the experiment?

4. What is the relationship between temperature and volume of the gas (assuming constant pressure)?

5. The pressure did not vary during the experiment, because all trials were performed at constant pressure. If the pressure had varied, how would it have affected the result?

6. Are the result consistent with the Charles's law which states that the volume of a given mass of gas is directly proportional to its Kelvin temperature at constant pressure? Explain.

7. In your interpretation of the results of this lab explain how the kinetic molecular theory of matter applies to the gases observed.

8. Does the result of the lab prove or disprove your hypothesis?

Conclusion: _____

_____

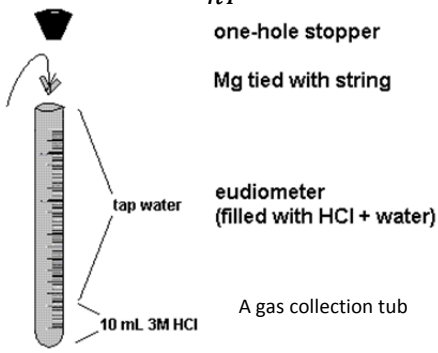
_____

_____

Ideal Gas Law Lab

Combining Avogadro's Law with the previous laws gives the ideal gas law, written pV = nRT. This is called an equation of state. Any gas (under ideal conditions – "relatively" high temperature and low pressure) must obey the ideal gas equation. Rearranging the ideal gas equation to solve for R, the ideal gas constant, gives R = $\frac{pV}{nT}$

Objective: In this lab, you will prepare a known amount of hydrogen gas at a known volume, temperature, and pressure, allowing you to substitute and calculate the value for R.



A gas collection tub

Procedure: You will be using a STRONG ACID. WEAR GOGGLES AND APRON AT ALL TIMES.

1. Fill the 1000 mL graduated cyclinder with water.
2. Measure the length of the magnesium ribbon given to you to the nearest millimeter.
3. Roll the magnesium into a loop and pass the thread through the loop.
4. Tilt the eudiometer and slowly add about 10 mL of 4M hydrochloric acid to the eudiometer tube.
5. Carefully and slowly fill the tube with water so the denser acid stays at the bottom of the tube.
6. Add the magnesium loop to the tube.
7. Cover the tube with your thumb making sure the thread end is held fast. Make sure there are no air bubbles.
8. Invert the eudiometer tube into the 1000 mL graduated cylinder. MAKDE SURE THE MOUTH IS COMPLETELY SUBMERGED INTO THE WATER BEFORE YOU RELEASE YOUR THUMB.
9. Release finger and hold tube so stopper stays underwater.
10. The acid will fall down to the magnesium and react with it to form hydrogen gas. The reaction is : $Mg + 2HCl \rightarrow H_2 + MgCl_2$. One mole of magnesium gives one mole of hydrogen gas with HCL in excess.
11. When the reaction is complete, equalize the water level inside and outside the tube to equalize the inside and outside pressures. Notice how the volume changes as you raise and lower the tube.
12. Note the pressure and temperature in the lab.
13. Clean up. Leave lab cleaner than you found it.

Data:

1	length of ribbon (cm)	
2	Linear density of Mg (g/cm)	0.53g/30 cm
3	Pressure (mm Hg)	
4	Temperature (°C)	
5	Vapor pressure (mm Hg)	
6	Volume of H_2 liberated (mL)	

Calculation (show work in space given)

7	Mass Mg (g)	
8	MM Mg (g/mol)	
9	Moles Mg (mol)	

10	Moles hydrogen gas	
11	Temperature (K)	
12	Volume of gas (L)	
13	Pressure of dry gas (mm Hg)	
14	R = $\frac{pV}{nT}$ R, $\frac{\text{L}\cdot\text{mm Hg}}{\text{mol}\cdot\text{K}}$	(experimental value)
15	Percent error Accepted value of R 62.4 $\frac{\text{L mm Hg}}{\text{mol K}}$	$\frac{\lvert \text{exptl value} - \text{accepted value} \rvert}{\text{accepted value}} \times 100$

Class Data: Compare your results with the other groups.

Group	**Experimental Value of R** $\frac{\text{L}\cdot\text{mm Hg}}{\text{mol}\cdot\text{K}}$	**Percent Error**
1		
2		
3		
4		
5		
6		
Average		

Conclusion

Compute the percent error for the class average of R compared to 62.4 $\frac{\text{L}\cdot\text{mm Hg}}{\text{mol}\cdot\text{K}}$ and discus how your group data compared.

Describe the ideal gas in terms of kinetic molecular theory.

How might a real gas differ from an ideal gas?

What are some possible sources of error in this experiment?

4.a Gaseous state is the simplest state of matter. Throughout our life we remain immersed in the ocean of **air** which is a mixture of gases. We spend our life in the lowermost layer of the atmosphere called troposphere, which is held to the surface of the earth by gravitational force. The thin layer of atmosphere is vital to our life. It shields us from harmful radiation. The most abundant gases in our atmosphere are dinitrogen _____ , dioxygen_____ , argon_____, and water vapor_____ .

List the eleven elements that exist as gases under normal conditions. Look at the balloons in your textbook periodic table.

Dalton's law of partial pressures Dalton found that each gas in a mixture exerts pressure independently of the other gases. **Dalton's law of partial pressures** states that the total pressure of a mixture of gases is equal to the sum of the pressures of all the gases in the mixture, as shown below.[5]

$P_{total} = P_1 + P_2 + P_3 + \ldots P_n$

The portion of the total pressure (P_{total}) exerted by one of the gases is called its partial pressure (P_n). The partial pressure of a gas depends on the number of moles of the gas, the size of the container, and the temperature of the mixture. The partial pressure of one mole of any gas is the same at a given temperature and pressure.

Example Problem *Finding the Partial Pressure of a Gas*

Air is made up of four main gases: N_2, O_2, Ar, and CO_2. Air pressure at sea level is approximately 760 mm Hg. Find the partial pressure of oxygen, given the following partial pressures: N_2, 594 mm Hg; Ar, 7.10 mm Hg; and CO_2, 0.27 mm Hg. Use Dalton's law of partial pressures to solve the problem.

$P_{total} = P_{N_2} + P_{Ar} + P_{CO_2} + P_{O_2}$

Practice Problems

1. What is the partial pressure of oxygen gas in a mixture of nitrogen gas and oxygen gas with a total pressure of 0.48 atm if the partial pressure of nitrogen gas is 0.24 atm?

2. Find the total pressure of a mixture that contains three gases with the following partial pressures: 6.6 kPa, 3.2 kPa, and 1.2 kPa.

3. Find the total pressure of a mixture that contains five gases with the following partial pressures: 7.81 kPa, 13.20 kPa, 2.43 kPa, 12.50 kPa, and 2500 Pa.

4. Find the partial pressure of ammonia in a mixture of three gases with a total pressure of 75.6 kPa if the sum of the partial pressures of the other two gases is 34.9

5. Nitrogen (80 kPa), oxygen (21.0 kPa), carbon dioxide (0.03 kPa), and water vapor (2.0 kPa) are the usual atmospheric components. What is the total atmospheric pressure in kPa?

---

[5] **4. i. *** Students know how to apply **Dalton's law of partial pressures** to describe the composition of gases and Graham's law to predict diffusion of gases based on temperature, pressure, and molar mass (size). It is important to distinguish clearly between diffusion and effusion. **Diffusion** is the process by which separate atoms or molecules intermingle as a result of random motion. **Effusion** is the process by which gas molecules pass from one container to another at lower pressure through a very small opening. Dalton's law of partial pressures states that total pressure in a gas-filled container is equal to the sum of the partial pressures of the component gases.

Gas Laws Review / Mole

1. Does 1 mole of a gas always occupy 22.4 liters? Explain

2. One mole of a diatomic gas is in a 22.4 liter flask at 0 °C.
   A. How many molecules of the diatomic gas are present in the flask?

   B. If the temperature is increased to room temperature, how many moles of the diatomic gas will be in the flask?

3 A. What effect does increasing temperature have on pressure?

   B. What effect does decreasing pressure have on temperature?

4. A. What effect does increasing pressure have on the volume of a gas?

   B. What effect does decreasing pressure have on the volume of a gas?

5. A. What effect does increasing temperature have on the volume of a gas?

   B. What effect does decreasing temperature have on the volume of a gas?

6. The <u>pressure on a gas is doubled</u> at constant temperature.
   a. Will the volume of the gas increase or decrease?

   b. By what factor will the volume change?

7. A 22.4 liter container contains 1 mole of gas at STP. Describe what would happen if the following changes were made on a system.
   a. Double the pressure by changing the volume.

   b. Double the absolute temperature.

8. Two glass containers have the same volume. One is filled with hydrogen gas, the other with carbon dioxide gas. Both containers are at the same temperature and pressure.
   a. Compare the number of moles of the two gases.

   b. Compare the number of grams of the two gases.

Chemistry: Practice Problems for the Gas Laws Do the following problems, showing work and all proper units.

1. A sample of gas has an initial volume of 25 L and an initial pressure of 3.5 atm. If the pressure changes to 1.3 atm, find the new volume, assuming that the temperature remains constant.

2. If a gas is cooled from 323.0 K to 273.15 K and the volume is kept constant what final pressure would result if the original pressure was 750.0 mm Hg?

3. If a gas in a closed container is pressurized from 15.0 atmospheres to 16.0 atmospheres and its original temperature was 25.0 °C, what would the final temperature of the gas be?

4. A sample of neon is at 89°C and 123 kPa. If the pressure changes to 145 kPa and the volume remains constant, find the new temperature, in °C.

5. A 30.0 L sample of nitrogen inside a rigid, metal container at 20.0 °C is placed inside an oven whose temperature is 50.0 °C. The pressure inside the container at 20.0 °C was at 3.00 atm. What is the pressure of the nitrogen after its temperature is increased?

9. A sample of sulfur dioxide (SO_2) is initially at a temperature of 133°C, a volume of 20 L, and a pressure of 850 mm Hg. If the volume changes to 25 L and the temperature increases to 181°C, find the new pressure.

Ideal Gas Law

10. 32 g of methane (CH_4) has a pressure of 450 kPa at 173°C. Find the volume occupied by the gas.

11. A sample of gas has a volume of 5.0 L when at a temperature of 320 K and a pressure of 2 atm.

    a) Find the number of moles of gas.

    b) If there are 15.2 g of the gas, calculate which noble gas is it? (Ar)

Gas Stoichiometry

1.a) Write a balanced chemical equation for the combustion of methane to form carbon dioxide and water.

b) If the methane has a volume of 0.65 L when under 100 kPa of pressure and at a temperature of 305 K, find the moles and then the mass of oxygen that is needed to use up all of the methane.

2.Boyle's Law states that the pressure of a gas is inversely proportional to its volume. Explain that statement. (Include the correct formula and examples)

3. A 7.0 liter balloon at room temperature (22°C) contains hydrogen gas. If the balloon is carried outside to where the temperature is –3.0°C, what volume will the balloon occupy?

4. A 5.0 liter tank of oxygen gas is at a pressure of 3 atm. What volume of oxygen will be available if the oxygen is used at standard pressure?

5. A 500 liter volume of helium gas is at a pressure of 750 mm Hg and has a temperature of 300K. What is the volume of the same gas at STP?

Chemistry Vocabulary: Gas Laws

Match each example below with the appropriate gas property it illustrates.

_____1. the fragrance of perfume spreads through the room	a. compressibility
_____2. smog forms over Sacramento during summer days	b. diffuses through other gases
_____3. a cylinder of oxygen used in a hospital	c. exerts pressure
_____4. shrink wrap	d. fills container
_____5. a balloon is inflated with helium	e. has mass
_____6. a balloon filled with air weighs more than an empty balloon	

Match the variables used to describe gases to the correct unit.

_____7. kPa	a. pressure
_____8. °C	b. temperature
_____9. mL	c. volume
_____10. K	
_____11. mm Hg	
_____12. atmospheres (atm)	
_____13. L	
_____14. °F	

Complete the following statements by writing "decreases," "increases," or "remains the same" on the line provided.

As a gas is compressed in a cylinder

15. its mass _____.
16. the number of gas molecules _____.
17. its pressure _____
18. its volume _____.
19. the distance between gas molecules _____.
20. its density _____.

Compete the following statements about the nature of gases as presented in the kinetic molecular theory by filling in the appropriate word (s) from the list below.

kinetic energy	no force	perfectly elastic	weak
potential energy	pressure	random motion	zero

25. Gas particles exert _____ on one another.
26. Gas molecules are said to be in _____.
27. The volume of gas particles themselves is said to be _____.
28. The collisions between gas particles are _____.
29. The temperature of a gas is a measure of the average _____ of the gas particles.

Kinetic Theory

- Gases are composed of _____, _____ particles called _____.
- Gas molecules are in _____ _____.
- All _____ between particles are _____ _____.
- The _____ of a gas display no _____ or _____ for one another.
- The _____ _____ _____ of the molecules is _____ _____ to the _____ temperature of the gas.

Ideal Gas

- Gas whose _____ conforms to the _____ _____—it is _____.

Gas Pressure: Pressure = _____

Atmospheric Pressure - the _____ the earth's _____ exerts due to its _____.

Barometer:

- Instrument used to measure _____ _____.
- Invented by _____

Normal Atmospheric Pressure

- Also called _____ _____
  _____
  _____
  _____
  _____

STP: _____ _____ and _____
_____
_____

Manometer: _____ used to measure _____ _____

- U-shaped tube _____ filled with _____
  One end _____ to _____ _____
  One end _____ to _____

1. The theory that explains the behavior of gases at the molecular level is called the _____ _____ which is based on assumptions about a theoretical gas often referred to as an _____ -_____.

2. Gases deviate most from ideal gas behavior under conditions of very low _____ and very high _____.

   --The molecules of an ideal gas display no _____ or _____ for one another.

   --Under ordinary conditions, an ideal gas consists chiefly of _____ space, which explains why gases are so easily compressed.

   -- Ideal gas particles travel in _____ lines until they collide with each other or with the walls of their container.

   --The collisions between the molecules of an ideal gas are completely _____.

   --The average kinetic energy of the molecules of an ideal gas is _____ proportional to the _____ temperature of the gas.

3. A gas exerts pressure on the walls of its container because gas molecules _____ with the walls of the container. So, the pressure exerted by a gas depends on two factors:

a)

b)

4. To measure gas pressure an instrument called a _____ is used.

5. The earth's atmosphere has weight, which creates _____ _____.

6. The instrument used to measure atmospheric pressure is the _____.

7. Standard Temperature and Pressure (or _____) is: _____ K _____ °C

_____ kPa _____ atm_____ mm Hg _____ torr

8. At 1 atm, the height of the _____ in a barometer is 760 mm.

9. Use the kinetic theory to explain why a helium filled balloon "shrinks" when it is taken from a warm room to the outside on a cold day.

_____

_____

_____

_____

_____

_____

_____

10. Use the kinetic theory to explain why bubble wrap pops when it is squeezed.

_____

_____

_____

_____

_____

_____

_____

11. Use the kinetic theory to explain why tire pressure increases when more air is added to a tire.

_____

_____

_____

_____

_____

_____

_____

Boyle's Law

✓ The _____ of a fixed _____ of gas varies _____ with the _____ at constant _____.

✓ _____

✓ _____

Kinetic Theory and Boyle's Law

✓ _____ of a gas is caused by the _____ of the gas _____ the walls of the _____.

✓ If the gas is _____ to _____ the volume it had, _____ as many _____ are present in any _____ _____.

* _____ as many _____ per _____ on the walls of the _____

* _____ of the gas will _____

Ex 1: A balloon filled with Helium has a volume of 457 mL at standard atmospheric pressure. After the balloon is released, it reaches an altitude of 6.3 km where the pressure is only 65.5 kPa. What is the volume of the balloon at this altitude?

Ex 2: Under a pressure of _____ mm Hg, a confined gas has a volume of _____mL. If the pressure is increased until the volume is _____ mL, what is the new pressure, assuming the temperature remains constant?

Charles's Law

✓ For a _____ _____ of gas, as long as the _____ is held _____, the _____ varies _____ with the _____ _____.

✓ _____

✓ _____

The Kelvin Temperature Scale

✓ _____ zero

* _____ possible _____

* _____ been reached

✓ _____ = absolute zero

✓ _____ = _____

✓ □K = _____

Ex 1: A quantity of gas occupies a volume of 506 cm^3 at a temperature of 147°C. Assuming the pressure stays constant, at what temperature will the volume of the gas be 604 cm^3?

Kinetic Molecular Theory and Charles's Law

✓ _____ the _____ _____ of a gas _____ the average _____ _____ of its _____.

✓ _____ moving molecules

* strike the walls of the _____ _____ _____

* strike the walls of the _____ with _____ _____

From _____ law we derive that the _____ would have to _____ if the _____ _____ is _____ so that _____ would remain _____.

Note Taking Guide: Episode 903 The Combined Gas Law

✓ Expresses the relationship between the _____, _____ and _____ of a _____ amount of _____.

_____ or _____

Ex: A sample of gas has a volume of _____ L when its temperature is _____ K and its pressure is _____ mm Hg. What volume will the gas occupy at STP?

V1 = _____ V2 = _____

T1 = _____ T2 = _____

P1 = _____ P2 = _____

Diffusion - The _____ spreading of a _____

Graham's Law of Diffusion

Under the same conditions of _____ and _____, gases _____ at a rate _____ proportional to the _____ _____ of their _____ (or _____ _____)

✓ _____ or _____

IDEAL GAS EQUATION: _____

New variables:

n = _____ of gas in _____

* _____ constant R = _____ _____ _____

* value depends on _____ used for _____ and _____

* value of R when using _____ and _____, R = _____

Ex: The average lung capacity for a female student is 3.9 L. At normal body temperature, 37°C, and 110 kPa, how many moles of air could her lungs hold?

P = _____ V = _____ T = _____

n = _____ R = _____

Avogadro's Law

Equal _____ of different _____ under the _____ conditions have the _____ number of _____.

Conversely, if samples of _____ _____ at the same _____ and _____ contain the _____ number of _____, then the _____ of all the _____ must be _____.

✓ At _____, one _____ of any gas occupies a _____ of _____.

✓ _____ is the _____ _____ of a gas.

Ex. 3.2 moles of KNO_3 are heated, producing O_2 and KNO_2. Calculate the volume of O_2 in liters, that could be obtained at STP.

Dalton's Law of Partial Pressures

✓ The _____ of a gas _____ is the _____ of the _____ _____ of each gas _____.

✓ _____

Ex: Oxygen gas has been collected over water at a total pressure of 95.0 kPa and a temperature of 25°C. What is the pressure of the dry oxygen gas?

Practice Questions SHOW ALL WORK

1. Gas pressure is caused by:

A. gas molecules heating up

B. gas molecules hitting the walls of a container

C. gas molecules hitting other gas molecules

D. gas molecules reacting with other gas molecules

2. "Absolute zero" is equal to:

a. 0 °C b. 0 °F c. 0 K d. 273 °C

3. No temperature can be reached that is below:

a. 0 Celsius b. 0 Kelvin c. 0 Fahrenheit d. 273 Kelvin

4. On the Kelvin scale, a temperature of 22 degrees Celsius has a value of:

a. 0 kelvins

b. 259 kelvins

c. 295 kelvins

d. -251 kelvins

5. Convert 300 °C to Kelvins

a. 300 °C = 1060 K

b. 300 °C = 573 K

c. 300 °C = -27 K

d. 300 °C = 27 K

6. On the Celsius scale, a temperature of 317 kelvins would have a value of:

a. 590 °Celsius b. 44 °Celsius c. 0 °Celsius d. 339 °Celsiu

7. Which pressures are equal to 2 atmospheres?

a. 760 torr b. 380 torr c. 152 kPa d. 1520 Torr

8. Which pairs represent Standard Temperature and Pressure (STP)?

a. . 273 °C and 760 atmospheres

b. 273 °C and 1 atmosphere

c. 0 °C and 760 torr

d. 0 K and 1 torr

9. At standard pressure, a sample of nitrogen occupies 500 mL. What volume does the gas occupy when the pressure doubles?

a. 250 mL

b. 2 mL

c. 1000 mL

d. 380 mL

10. At a pressure of 5.0 atmospheres, a sample of gas occupies 40. liters. What volume will the same sample occupy at 1.0 atmosphere.

a. 8.0 liters

b. 0.0050 liters

c. 200 liters

d. 0.13 liters

11. At constant pressure and 25 °C a sample of gas occupies 4.5 L. At what temperature will the gas occupy 9.0 L?

a. 596 K

b. 50 °C

c. 596 °C

d. 50 K

12. A small sample of helium gas occupies 6 mL at a temperature of 250 K. At what temperature does the volume expand to 9 mL?
a. 500 K
b. 125 K
c. 375 K
d. 2250 K

13. In a closed container at 1.0 atmosphere, the temperature of a sample of gas is raised from 300 K to 400 K. What will be the final pressure of the gas?
a. 100 atmospheres
b. 0.010 atmospheres
c. 0 atmospheres
d. 1.3 atmospheres

14. Organize the following gases in order of their rates of diffusion, from fastest to slowest: O_2, NH_3, H_2, CO_2,
a. hydrogen, oxygen, ammonia, carbon dioxide
b. hydrogen ,carbon dioxide, oxygen, ammonia,
c. hydrogen, ammonia, oxygen, carbon dioxide
d. hydrogen, oxygen, carbon dioxide, ammonia

15. A sample of H_2 gas is held in a 1-L metal cylinder. At what temperature will the gas exert the most pressure?
a. 25 °C b. 55 °C c. 35 °C d. 40 °C

16.Which of the following is defined as a measure of the average kinetic energy of particles in a given sample?
A Velocity B Diffusion C Temperature D Partial pressure

17 Energy release is to condensation as energy input is to —
A deposition B sublimation C freezing D dispersion

18 Which of the following is <u>NOT</u> a characteristic of liquids?
A No significant attraction between particles
B Less fluid than gases
C More dense than gases
D Exhibits viscosity

19 Marta and her father often skip stones across a pond. What type of intermolecular force creates the surface tension that allows the stones to skip?
A. Metallic forces
B. Dipole–dipole forces
C. Dispersion forces
D. Hydrogen bonding

20 During evaporation, certain liquid molecules become vapor molecules because they have greater than average —
A. lattice energy B. viscosity C. kinetic energy D. fluidity

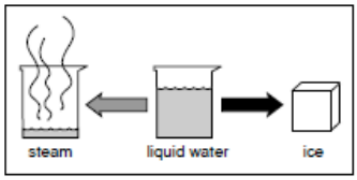
21. Which of the following is a gas–gas behavior relationship?
A. Helium gas is heated and its volume increases.
B. Oxygen gas is compressed and its temperature increases.
C. Nitrogen gas is placed in a container and the molecules settle to the bottom.
D. Hydrogen gas is cooled and its pressure increases.

22. Which of these decreases as a given volume of gas increases?
A Number of gas particles
B Temperature
C Pressure
D Kinetic energy

23 Ionic solids such as sodium chloride are easily shattered, but metallic solids such as copper can be easily bent and shaped. This difference occurs because —
A ionic solids have low melting points
B atoms in metallic solids are not arranged in a regular pattern
C covalent bonding between sodium and chlorine keeps the solid rigid
D mobile electrons in the copper can shift without disrupting the solid

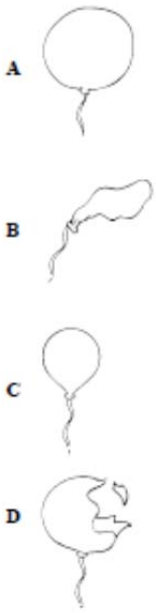
24 Diffusion is the term used to describe the movement of one material through another. The diffusion of gases can be explained by —
A relative molar masses
B differences in volume
C evaporation
D random motion

25. The diagram shows how liquid water is transformed into a solid and a vapor.
Which label should be placed above each of the arrows in the diagram?



a. Energy added over the gray arrow; energy released over the black arrow
b. Particle velocity decreased over the gray arrow; particle velocity increased over the black arrow
c. Energy released over the gray arrow; energy added over the black arrow
d. Density decreased over the gray arrow; density increased over the black arrow

26. The kinetic-molecular theory of gases explains the behavior of gases at the molecular level.
All of the following are part of this theory EXCEPT —
A gas molecules experience completely elastic collisions
B all gas molecules have the same average kinetic energy at the same temperature
C gas particles are in constant, random motion
D gas molecules are incompressible



27 You are given a balloon filled with a known volume of helium gas. You place the balloon inside a freezer for an hour. How will the balloon look after being in the freezer?

28 Physicians can use liquid nitrogen to freeze and destroy warts and other skin growths.
Knowing the assumptions of the universal gas law, this should surprise you most because —
A. if a gas can liquefy, that would imply that gases experience intermolecular forces
B. all gases are volatile and can't be used indoors
C. gas particles are too small to be condensed
D. if a gas can freeze, that would imply that gases can be kept at cold temperatures

29 David has two containers of two different gases at the same temperature and pressure. David could assume all of following EXCEPT —
A. when the temperature is increased, the volume of both containers will increase
B. when the pressure is increased, the volume of both containers will decrease
C. both containers contain the same number of gas particles
D. when the pressure is decreased, the temperature of both containers will increase

30. Air bags, which act as safety devices in cars, contain solid sodium azide. On impact, the sodium azide releases nitrogen gas, which expands the air bag. The main benefit of using a gas instead of another type is that —
A. gas molecules are subject to ionic bonding
B. the separation of gas molecules is much greater than the volume they occupy
C. gases won't explode the bag on very hot days
D. gas molecules don't transfer excess Kinetic Energy

Bellringers

1/6/14 Solve the following conversion problems Show all work

7.35 psi to atm

1520Torr to kPa

-23 °C to K

498 K to °C

1/7/14

Convert 25 mL to L

At STP 1 mole of a gas occupies how many mLs

STP stands for _____ _____ and _____ which is:

1/8/14

List the five tenets of the Kinetic Molecular Theory

_____

_____

_____

_____

_____

1/9/14

The 3 particles and respective charges of the atom are:

a. _____ b. _____ c. _____

The number of protons in one atom of an element determines the atom's_____ , and the number of electrons determines _____ of the element.

Name the element which has the following numbers of particles:

a. 26 electrons, 29 neutrons, 26 protons _____ b. 53 protons, 74 neutrons _____

c. 2 electrons (neutral atoms) _____ d. 0 neutrons _____

e. 86 electrons, 125 neutrons, 82 protons _____

1/10/14 Select the elements in order of increasing ionization energy.

Arrange He, K, Fr in order of increasing ionization energy, then select the correct answer

a. He, K, Fr c. Fr, K, He

b. K, Fr, He d. Fr, He, K

.

Describe the periodic trend for ionization energy (refer to group trends and period trends) and explain ionization energy

_____

_____

_____

1/13/14 Absolute zero is 0K and is the temperature at which molecular motion by kinetic energy stops. What is this temperature in °C?

1/14/14 Three moles of carbon dioxide are produced when one mole of propane gas C_3H_8 is burned. How many moles of carbon dioxide will be produced if 30 moles of propane gas are burned?

A 10 moles

B 30 moles

C 90 moles

D 120 moles

1/15/14

In the movie *The Wacky World of Chemistry,* a chemist wrote down the following equation on a chalkboard:

$Ti + C + 2Cl_2 \rightarrow TiCl_3 + C$. This equation is NOT correct because —

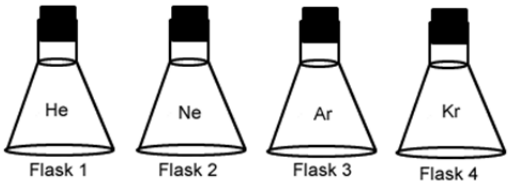
A the titanium atoms are not equal on both sides of the equation

B there are not enough chlorine atoms on the right side of the equation

C the carbon atoms are equal on both sides of the equation

D the right side of the equation should have a greater number of atoms than the left side

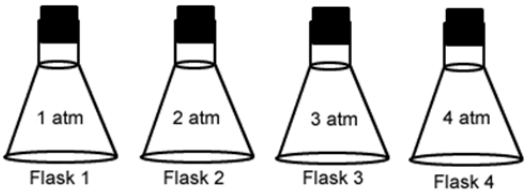
1/16/14



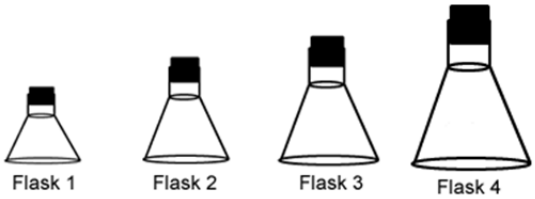
If all of the following flasks are the same size, at the same temperature, and contain the same number of molecules, in which flask will the molecules be moving fastest?

Why?_____

Each of these flasks is the same size and at the same temperature. Which one contains the most molecules? Why?



_____



Each of these flasks contains the same number of molecules. In which container is the pressure highest?_____

1/21/14 A 200 mL sample of H_2 gas is collected at a pressure of 1.00 atm. If the temperature remains constant, what volume will the gas occupy at 380 mm Hg? Convert pressures to same scale!

An average adult has a "deep breath" lung capacity of about 4.5 L. 4.50 L of air is measured at 740 mm Hg pressure. It is compressed until its volume is 3.50 L. What is its new pressure (assuming unchanged temperature)?

1/22/14 Some CO_2 gas at 760 mm Hg pressure and – 73.0°C in a sealed metal container is then heated to 27.0° C. What is its new pressure?

1/23/14 The pressure in an automobile tire is 2.00 atm at 27.0°C. At the end of a road trip the pressure has risen to 2.20 atm. What is the temperature of the air in the tire, assuming the volume has not changed? Calculate the final T in Kelvin, then change to °C.

1/24/14 A balloon is placed in a freezer and has a volume of 500 ml at – 23.0°C. It is then removed from the freezer. What will be its volume at + 27.0°C (assuming constant pressure)?

A sample of nitrogen occupies a pressure of 250kPa at 250 K. What pressure will it be at 125K?

1/27 What would happen to the *volume* of a gas under the following conditions?
(gets smaller, gets larger, stays the same, or can't tell)

a. Increase Temp, keep pressure and amount constant _____

b. Decrease pressure, keep temp and amount constant _____

c. Increase pressure, add gas, temp constant _____

d. Increase temp, decrease pressure, amount constant _____

What would happen to the pressure and kinetic energy of a gas in an enclosed metal container when

e. Add gas to a metal cylinder _____

f. Double the Kelvin temp_____

GIVEN THE COMBINED GAS LAW EQUATION ON THE BACK OF YOUR PERIODIC TABLE:
REARRANGE TO DERIVE:

P_1

T_2

V_1

T_1

V_2