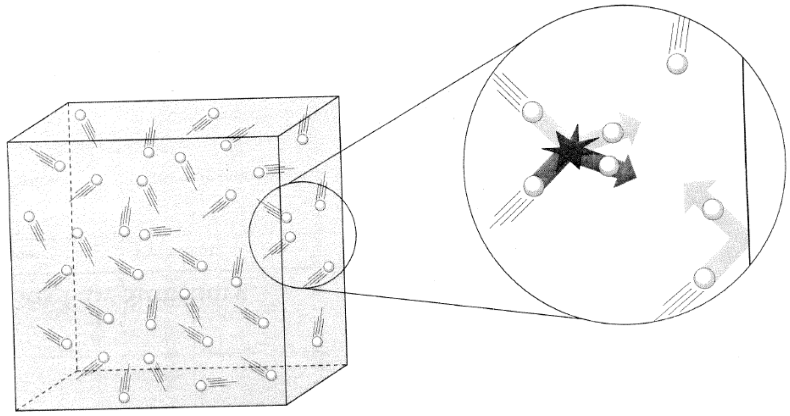
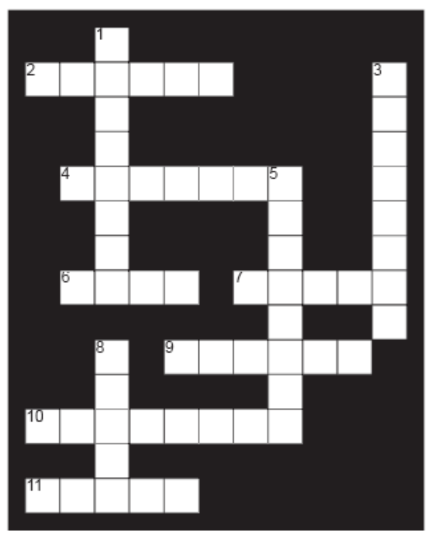
Gases – Unit 2



Element Crossword Puzzles

It's a Gas!



Across

2. This gas can be used to fill party balloons.
4. This gas is used to make yellow-green 'neon' signs.
6. This gas is widely used in luminous signs.
7. Nearly 1% of the earth's atmosphere is...
9. About 21% of the earth's atmosphere is...
10. Most of the visible universe is made of...
11. This gas is used in strobe lights.

Down

1. When combined with tin, this gas helps keep your teeth strong.
3. This gas helps keep swimming pools clean.
5. About 78% of the earth's atmosphere is...
8. A radioactive gas.

What we already know is that….	**Examples we know of are….**
What we have heard in the media…	**Gases are like ____________because…**

GASE

Explaining Boyle’s Law

Power Point slides 14-28

Introduction

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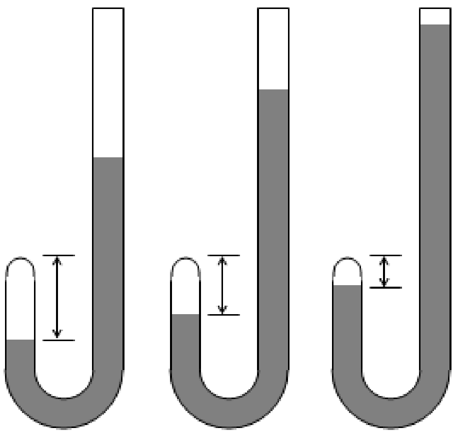
Robert Boyle’s Experiment

The experiment that Boyle used to study the relationship between pressure and volume involved filling a closed-ended tube, called a J-tube, with mercury, trapping gases in the closed end.



Results:

There is no such thing as “sucking”.

- A practical application of Boyle’s Law is drinking through a straw.

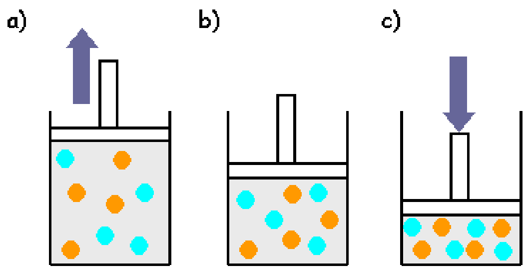
Diagram

Explaining Boyle’s Law:

How could we use the kinetic molecular theory to explain how Boyle's Law works?

Answer:

Basically:



Let's try a Boyle's Law game (more commonly known as Jezzball) - http://mokun.tripod.com/java/jezzballp.htm

How does this game accurately represent gas molecules?
1.
2.
3.

Practical Applications of Boyle's Law - Breathing

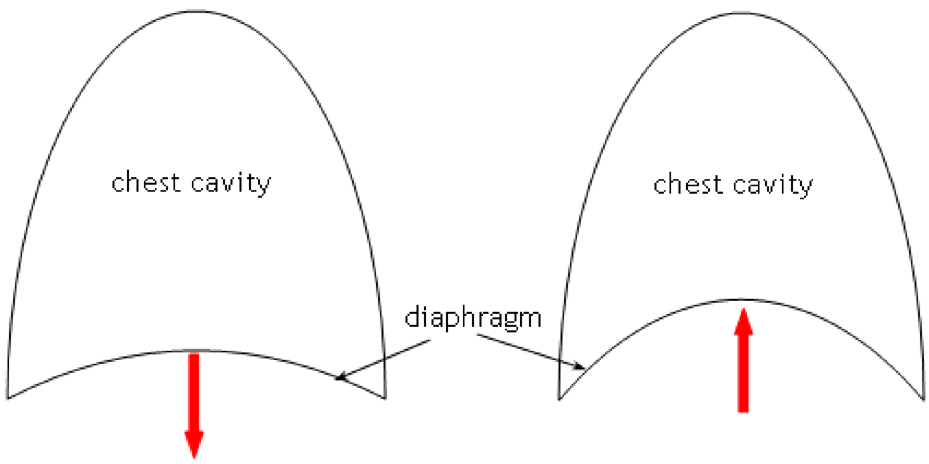


Figure 1. Inhale Figure 2. Exhale

Inhaling:

- The diaphragm moves downward allowing the lungs to increase in volume. (Fig 1)
- The increased lung volume decreases the pressure in the chest cavity so that it is less than the air pressure.
- The lower pressure forces air to rush into the lungs to equalize the pressure in the chest cavity. (Gas diffusion)

Exhaling:

- When we exhale the diaphragm moves upward decreasing the volume of the chest cavity.
- The decreased volume increases the pressure in the lungs until the pressure in the lungs is greater than the air pressure.
- The increased pressure forces the air out of the lungs.

Challenge – Can you pull the bag out of the beaker?
Boyle's Law Equation:

Example 1
If 3 L of gas is initially at a pressure of 1 atm, what would be the new pressure to cause the volume of the gas to become 0.5 L?

Example 2.
A syringe contains 20 mL of a gas at 100 kPa. The pressure in the syringe is changed to 25 kPa. What is the new volume of the gas?

Pressure – Volume Exercises

Answer the following questions. Remember that showing all your work is good practice.

1. Gas is placed into a syringe until the pressure is 45.0 kPa. What is the new pressure if

   a) the volume in the syringe is doubled?

$P_1V_1 = P_2V_2$

$(45)(1) = P_2\ (2)$

$P_2 = 22.5$ kPa

   b) the volume in the syringe is tripled?

$P_1V_1 = P_2V_2$

$(45)(1) = P_2(3)$

$P_2 = 15$ kPa

c) the volume is one third its original volume?

$P_1V_1 = P_2V_2$

$(45)(3) = P_2(1)$

$P_2 = 135$ kPa

2. 100.0 mL of gas is placed into a syringe. What is the new volume if

a) the pressure is doubled?

$P_1V_1 = P_2V_2$

$(1)\ (100) = (2)\ V_2$

$V_2 = 50$ mL

b) the pressure is tripled?

$P_1V_1 = P_2V_2$

$(1)\ (100) = (3)\ V_2$

$V_2 = 33.3$ mL

c) the pressure is one quarter the original pressure?

$P_1V_1 = P_2V_2$

$(4)\ (100) = (1)\ V_2$

$V_2 = 400$ mL

3. Change the following from the initial conditions to the new conditions:
   a) 100.0 mL oxygen gas at 10.50 kPa is changed to 9.91 kPa

$P_1V_1 = P_2V_2$

$(100)\ (10.5) = (9.91)\ V_2$

$V_{2\ =}$ 106.0 mL

   b) 50.0 cm^3 helium at 97.3 kPa is changed to 102.5 kPa

$P_1V_1 = P_2V_2$

$(97.3)(50.0) = (102.5)\ V_2$

$V_2 = 47.5$

   c) 25.0 mL nitrogen at 0.990 atm is changed to 0.751 atm

$P_1V_1 = P_2V_2$

$(0.990)\ (25.0) = (0.751)\ V_2$

$V_2 = 33.0$ mL

   d) 745 torr of hydrogen in 0.550 L is changed to 0.700 L

$P_1V_1 = P_2V_2$

$(745)\ (0.550) = P_2\ (0.770)$

$P_2 = 532$ torr

   e) 1.40 atm of carbon dioxide in 1.32 L is changed to 0.705 L

$P_1V_1 = P_2V_2$

$(1.40)\ (1.32) = P_2\ (0.705)$

$P_2 = 2.62$ atm

f) 525 mL neon at 49.3 kPa is changed to 845 mL

$P_1V_1 = P_2V_2$

$(49.3)(525) = P_2\ (845)$

$P_2 = 30.6$ kPa

f) 0.150 L carbon monoxide at 635 mmHg is changed to 895 mmHg

$P_1V_1 = P_2V_2$

$(635)(0.150) = (895)\ V_2$

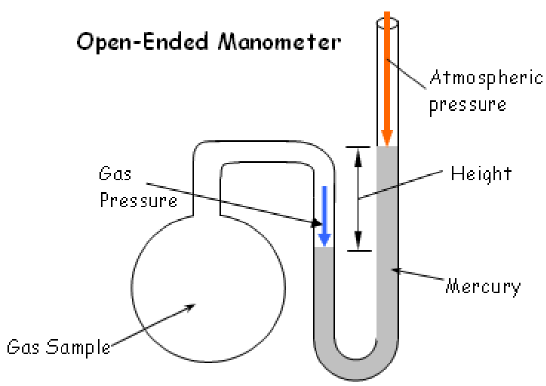
$V_2 = 0.106$ L

Measuring Gas Pressure in mmHg

Pressure can be measured using several devices or instruments. Air pressure or atmospheric pressure is measured with a barometer, whereas gas pressure is measured with a manometer.

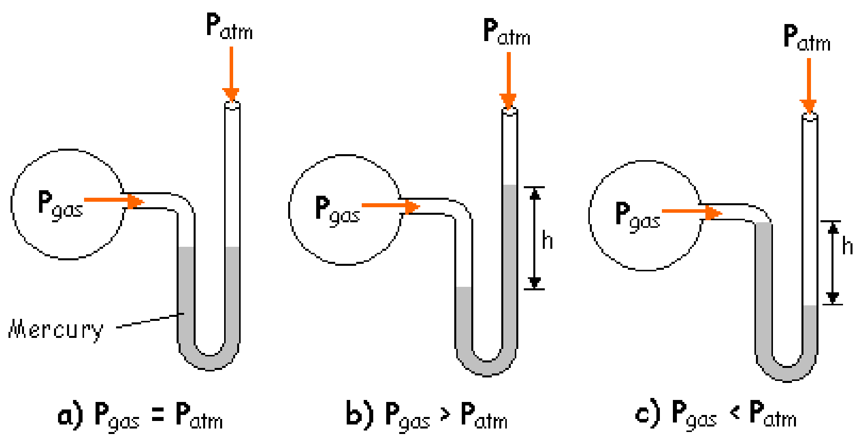
A **manometer** usually has a bulb or glass container on one end and can be open or closed on the other. A liquid, often mercury, is placed in a U-shaped tube.

The pressure is measured by finding the difference in height on both sides of the tube (see figure below). In the open-ended manometer, the pressure of the gas is related to the height difference, *h*, of the mercury (or other liquid) in both sides of the U-tube. It is called open-ended because one end is exposed to the gas pressure and the other end is open to the atmosphere.



When the pressure of the trapped gas (Pgas) is equal to the atmospheric pressure (Patm) the height on both sides of the U-tube are equal (a in the figure below).

When the pressure of the trapped gas is greater than the atmospheric pressure, the height of the liquid in the open side of the U-tube will be greater than the closed side (b in the figure below). The opposite is true if the atmospheric pressure is greater than the gas pressure (c in the figure below).



To calculate the pressure of the trapped gas in (b), when the gas pressure is greater than the atmospheric pressure (Pgas > Patm)

$$\mathbf{Pgas = Patm + } \boldsymbol{h}$$

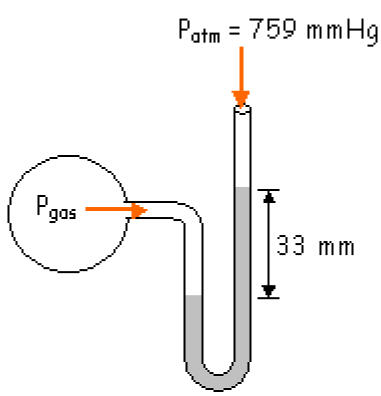
In (c) above, when atmospheric pressure is greater than gas pressure (Pgas < Patm)

$$\mathbf{Pgas = Patm - } \boldsymbol{h}$$

Gas Questions

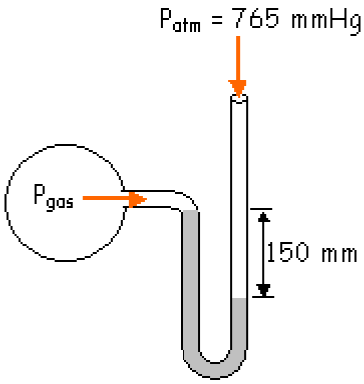
1. Find the pressure in mm Hg of the gas in each of the following manometers:

   a.



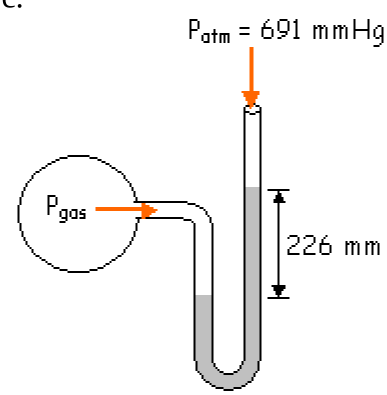
$P_{gas} = P_{atm} + h$
$P_{gas} = 759 + 33$
$P_{gas} = 792$ mmHg

b.



$P_{gas} = P_{atm} - h$
$P_{gas} = 765 - 150$
$P_{gas} = 615$ mmHg

c.



$P_{gas} = P_{atm} + h$
$P_{gas} = 691 + 226$
$P_{gas} = 917$ mmHg

2. A gas container is fitted with a manometer. The level of the mercury is 15.0 mm lower on the open side. Using a laboratory barometer, you find that atmospheric pressure is 750.0 mm Hg. What is the pressure of the gas in the container?

$P_{gas} = P_{atm} - h$
$P_{gas} = 750 - 15$
$P_{gas} = 735$ mmHg

3. A soccer ball is attached to an open-ended manometer. The mercury level in the manometer is 10.0 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 770.0 mm Hg. What is the gas pressure, in mmHg, inside the ball?

   $P_{gas} = P_{atm} - h$
   $P_{gas = 770 - 10}$
   $P_{gas = 760 \text{ mmHg}}$

Try Practice Problems 13-2 #9-15 (page 11 of this booklet)

Converting Units of Pressure

Gas Pressure is the outwards force (push) a gas has on another object or surface. If there is a lot of force in a small area then the pressure is high. If there is less force in a larger area then the pressure is low. Think about the force a high heeled shoe has on the floor as opposed to a flat running shoe.

There are four basic units of measurement for pressure:

1. The unit **atmosphere** (**atm**) was derived from standard atmospheric pressure at sea level. Two atmospheres is twice standard atmospheric pressure, and so on.
2. The SI unit for pressure is the **Pascal** (**Pa**), named after Blaise Pascal. The Pascal is defined as 1 Newton of force per square metre. This is rather small unit, so when measuring gas pressure we use **kiloPascals**, which is 1000 Pascals.
3. **Millimetres of mercury** (**mmHg**) is a common unit used inside the laboratory today, however, many barometers found in the home use both mm of mercury as well as another unit like kiloPascals.
4. The **millibar** is a meteorological unit of atmospheric pressure. One bar is equal to standard atmospheric pressure or 1 atmosphere.
5. One other common pressure unit is **pounds per square inch** (**psi**). This is an Imperial unit of pressure. This unit is often encountered when filling car and bicycle tires.

Conversions:

- 1 atmosphere is equal to 760 mmHg, 14.7 psi or 101.325 kPa.
- (1atm = 760mmHg = 14.7psi = 101.325 kPa)

Example using length –

Questions:

1. If the atmospheric pressure is 1.5 atm what is the pressure in kPa?

$\frac{1.5\ atm}{1\ atm} = \frac{x}{101.325\ kPa}$ x = 151.988 kPa

2. If a gas pressure is recorded at 380 mmHg what would that pressure be in atm?

380 $\frac{mmHg}{760\ mmHg} = \frac{x}{1\ atm}$ x = 0.5 atm

3. If a vapour pressure is listed at 49.4psi what is that pressure in mmHg?

$\frac{49.4\ psi}{14.7\ psi} = \frac{x}{760\ mmHg}$ x = 2554.0 mmHg

Try Practice Problems 13-2 #1-8 (page 11 of this booklet)

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?

Answers: #1 – 1.07 atm, #2 – 288.8 mmHg, #3 – 813.2 mmHg, #4 – 0.998 atm, #5 – 101.498 kPa, #6 – 1.678 kPa, #7 – 2.939, #8 – 239.871, #9 – 735 mmHg, #10 – 760 mmHg, #11 – 759.8 mmHg, #12 - 766 mmHg = 102.125 kPa, #13 – 264 mmHg, #14 – 675.104 mmHg = 90.006 kPa, #15 – 801.4 mmHg

Boyle's Law Combined Questions – Great Practice!

1. If a soccer ball's volume changes from 4.2L to 265 mL after a plane ride, and the initial pressure was 2.5 atm, what is the new pressure in kPa?

$P_1V_1 = P_2V_2$
(2.5 atm)(4.2 L) = P_2 (0.265L)
P_2 = 39.623 atm

$$\frac{39.6\ atm}{1\ atm} = \frac{x}{101.325\ kPa}$$

x = 4014.764 kPa

2. A volleyball initially at 14.5 psi has its pressure altered to 1.5 atm. The initial volume of the volleyball was 3.2L, what is its new volume?

$$\frac{14.5\ psi}{14.7\ psi} = \frac{x}{1\ atm}$$

X = 0.986 atm

$P_1V_1 = P_2V_2$
(0.986 atm)(3.2 L) = (1.5 atm)V_2
V_2 = 2.1 L

3. A reaction between hydrogen and oxygen yields 250mL of water vapour. The pressure is 95 kPa. The gas is transferred to a 1.2L collecting tube, what is its new pressure?

$P_1V_1 = P_2V_2$
(95 kPa)(0.250) = P_2(1.2 L)
P_2 = 19.791 kPa

4. A balloon store traditionally fills balloons to 4.5L and a pressure of 845mmHg. They are now trying a new type of balloon that keeps the helium at a pressure of 234 KPa. How has this affected the volume of the balloon?

$$\frac{845\ \text{mmHg}}{760\ mmHg} = \frac{x}{101.325\ kPa}$$

x = 112.657 kPa

$P_1V_1 = P_2V_2$
(112.657)(4.5) = (234)V_2
V_2 = 2.166 L

5. The Colorado Avalanche have oxygen tanks behind their bench during games to help them breathe. The Tanks are 52L and hold the air at a pressure of 1.2 atm. They have just purchased smaller tanks (45L) that hold the same amount of air but under a greater pressure. What is the new pressure of their air tanks in psi?

$P_1V_1 = P_2V_2$
$(1.2)(52) = P_2\ (45)$
$1.387\ \text{atm} = P_2$

$$\frac{1.387\ atm}{1\ atm} = \frac{x}{14.7\ psi}$$

X = 20.384 psi

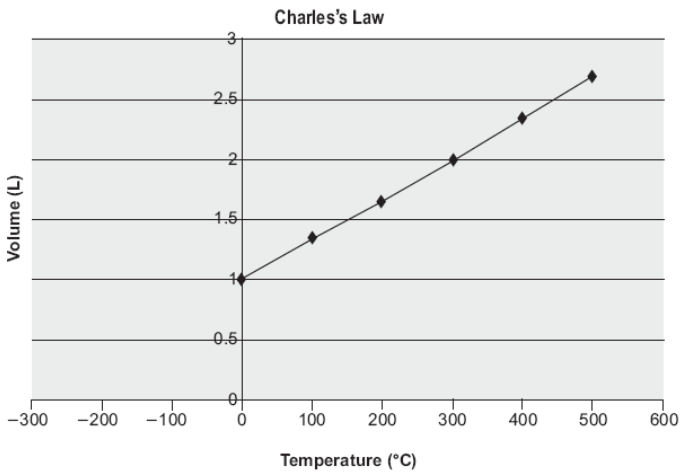
Gas Laws - Charles's Law

Charles Law simulation:
http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/flashfiles/gaslaw/charles_law.html

http://www.grc.nasa.gov/WWW/k-12/airplane/aglussac.html

The next significant advance in the study of gases came in the early 1800's in France. Hot air balloons were extremely popular at that time and scientists were eager to improve the performance of their balloons. Two of the prominent French scientists, **Jacques Charles** and **Joseph-Louis Gay-Lussac**, made detailed measurements on how the volume of a gas was affected by the temperature of the gas. Given the interest in hot air balloon at that time, it is easy to understand why these men should be interested in the temperature-volume relationship for a gas.

Just as Robert Boyle made efforts to keep all properties of the gas constant except for the pressure and volume, so Jacques Charles took care to keep all properties of the gas constant except for temperature and volume. The equipment used by Jacques Charles was very similar to that employed by Robert Boyle. A quantity of gas was trapped in a J-shaped glass tube that was sealed at one end. This tube was immersed in a water bath; by changing the temperature of the water, Charles was able to change the temperature of the gas.



Absolute Zero

Absolute zero is the temperature where gas volume drops to zero. A negative volume of gas is essentially impossible so this absolute temperature must be the lowest temperature that is possible.

The Kelvin scale was then developed based on this absolute zero. It was determined by William Thomson that –273°C is the lowest temperature possible, or absolute zero. The advantage of this scale is that there are no negative numbers.

$$K = °C + 273.15$$

Go back and complete the chart from your lab by converting from degrees Celsius to degrees Kelvin.

Example:

	Kelvin Scale	Celsius Scale
Absolute zero	0 K	–273.15°C
Freezing point of water	273.15 K	0°C
Boiling point of water	373.15 K	100°C

Sample Problems

We now solve Charles's Law by converting to Kelvins and by using the following equation. Also make a hypothesis about what will happen to the volume based on the temperature change to check your answers.

$$\underline{V_1} = \underline{V_2}$$

T_1 T_2

Problem 1:
If the temperature of 6.00 L of a gas at 25.0°C is increased to 227°C, determine the volume at the new temperature.

Solution:
First change the temperatures to kelvins.
25.0°C + 273 = 298 K
227°C + 273 = 500 K

If the temperature increases what happens to the volume?
It increases.

V2 = (V1 X T2) / T1
Volume = (6.00L x 500 K) / 298 K
Volume = 10.1 L

This is an increase in volume which matches our prediction.

Problem 2:
If the volume of a gas at –73.0°C is doubled to 48.0 L, calculate the final temperature in Celsius.

Solution:

Temperature-Volume Relationship Exercises
Description: Charles' Law
Remember to check the answer key binder.

Questions:

1. Calculate the change in temperature, in degrees Celsius, when the volume of 1.0L of a gas at 25°C is doubled to 2.0L.

273.15 + 25 = 298.15 K

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

$$\frac{1.0\ L}{298.15\ K} = \frac{2.0L}{T_2}$$

$T_2 =$ 596.3 K

596.3 – 273.15 = 323.15 °C

2. Calculate the new volume of 100 mL gas if its temperature is doubled from 20^{o}C to 40^{o}C.

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

$$\frac{100\ mL}{293.15\ K} = \frac{V_2}{313.15\ K}$$

V_2 = 106.822 mL

3. You get a balloon at the circus with a volume of 2.5 L at room temperature (25^{o}C). When you step outside it is winter and the temperature is -25^{o}C. Assuming pressure remains constant what is the balloons new volume?

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

$$\frac{2.5\ L}{298.15\ K} = \frac{V_2}{248.15\ K}$$

V_2 = 2.08 L

4. 20.0 mL of a gas at 285 K increases in volume to 32.0 mL. What is the new temperature of the gas?

$$\frac{20\ mL}{285\ K} = \frac{32\ mL}{T_2}$$

5. Find the new volume if the following changes occur at a constant pressure:
   a. 225 mL of oxygen at 273 K warmed to 300 K.
   b. 3.0 L of nitrogen gas is cooled from 90.0^{o}C to -45.0^{o}C.

a)

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

$$\frac{225}{273} = \frac{V_2}{300}$$

$V_2 = 247.2$ mL

b)

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

$$\frac{3.0L}{363.15\ K} = \frac{V_2}{228.15\ K}$$

$V_2 = 1.884$

6. Find the new temperature in degrees Celsius when the following changes occur at a constant pressure.
    a. 5.0 L of air at 45°C expands to 22.0L.
    b. 100.0 mL of Helium at 300 K compresses to 80.0 mL.

a)

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

$$\frac{5.0L}{318.15\ K} = \frac{22.0L}{T_2}$$

$T_2 = 1399.86$ K

1399.86 – 273.15 = 1126.71 °C

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

$$\frac{100\ mL}{300\ K} = \frac{80\ mL}{T_2}$$

$T_2 = 240\ K$

240 – 273.15 = -33.15 °C

Keep going:

1. If the relationship between temperature and volume of a gas at a constant pressure is linear, explain why the volume of a gas doesn't double when its temperature doubles from 10.0°C to 20.0°C.

The temperature isn't doubling from 10°C to 20°C. You can only think of temperature doubling in Kelvins.

2. A hot air balloon has a volume of about 2.5×10^6 L. The balloon is filled with 1.9×10^6 L of air at 18.0°C using a fan, then the propane heater is turned on. What temperature must the air in the balloon reach to fill the balloon?

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

$$\frac{1.9 \times 10^6}{291.15K} = \frac{2.5 \times 10^6}{T_2}$$

T_2 = 383.09 K
383.09 – 273.15 = 109.9 °C

3. A sample of gas at 20.0°C was heated to 100.0°C. If the volume of the gas was 250.0 mL after heating, what was its original volume?

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

$$\frac{V_1}{293.15K} = \frac{250\ mL}{373.15\ K}$$

V_1 = 196.4 mL

Some light reading…….

Absolute Hot

By Peter Tyson
Posted 01.08.08
NOVA

Is there an opposite of absolute zero, the lowest possible temperature?

The question *Is there an opposite to absolute zero?* seems innocent enough, right? Absolute zero is 0 on the Kelvin scale, or about minus 460°F. You can't get colder than that; it would be like trying to go south from the South Pole. Is there a corresponding maximum possible temperature?

Well, the answer, depending on which theoretical physicist you ask, is yes, no, or maybe.

For many of us, the hottest thing we could think of might be the core of the sun. But broiling as it is at roughly 10^7 degrees, it's a full 25 orders of magnitude colder than the current highest temperature that physicists propose.

Huh? you ask. Yeah, that's how I felt. And the question doesn't just mess with the minds of physics dummies like me. Several physicists begged off of trying to answer it, referring me to colleagues. Even ones who did talk about it said things like "It's a little bit out of my comfort zone" and "I think I'd like to ruminate over it." After I posed it to one cosmologist, there was dead silence on the other end of the line for long enough that I wondered if we had a dropped call.

Contender #1—10^{32} K

It's called the Planck temperature, after the German physicist Max Planck, and it equals about 100 million million million million million degrees, or 10^{32} Kelvin. "It's ridiculous is what it is," said Columbia physicist Arlin Crotts when I asked him if he could please put that number in perspective for me. "It's a billion billion times the largest temperature that we have to think about" (in gamma-ray bursts and quasars, for instance). Oh, that helped.

Truthfully, when contemplating the Planck temperature, you can forget perspective. All the usual terms for very hot—scorching, broiling, hellish, insert your favorite here—prove ludicrously inadequate. In short, saying 10^{32} K is hot is like saying the universe occupies some space.

In conventional physics—that is, the kind that relies on Einstein's theory of general relativity to describe the very large and quantum mechanics to describe the very small—the Planck temperature was reached 10^{-43} seconds after the Big Bang got under way. At that instant, known as one Planck time, the entire universe is thought to have been the Planck length, or 10^{-35} meters. *(More on Plank in grade 12 chemistry.)*

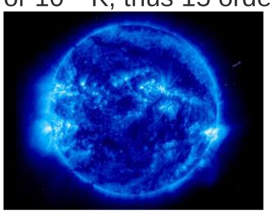
Contender #2—10^{30} K

String theorists, those physicists who believe the universe at its most fundamental consists not of particles but of tiny, vibrating strings, have their own take on temperature. This model posits a maximum temperature called the Hagedorn temperature. While string theorists don't give a specific number for the Hagedorn temperature, scientists have reasons to think it's about one percent of its theoretical cousin, the Planck. That makes it about 10^{30} K, or two orders of magnitude below the Planck.

Contender #3—10^{17} K

I learned of yet another highest possible temperature from an assistant professor of physics at Penn State, who is eagerly awaiting the day that officials at CERN on the Swiss-French border switch on the Large Hadron Collider, the world's largest particle accelerator.

One reason why they're excited has to do with temperature. As the researcher told me, "It may be that the [highest possible] temperature is—as I believe—the temperature or the energy right around the energy that the LHC will be probing." The LHC will operate at 14 trillion electron volts, or 10^{17} K, thus 15 orders of magnitude below the Planck.



Could absolute cold and absolute hot—whatever it is, if it even is—be manifestations of the same physical phenomenon? Here, an ultraviolet image of the sun's corona.

Contender #4—0 K

As if at least three different possible opposites to absolute zero weren't pause-giving enough, what Alexander told me next really set my head spinning. Whatever the highest temperature is, he said, it might, just might be essentially equivalent to the coldest temperature. "In other words, zero temperature is the same, in a sense, as the Planck temperature."

Alexander described two potential ways the universe began. Either it was at the Planck temperature and then inflated and cooled to create what we see today. Or it started off at zero temperature and speeded up as it expanded. "So one of two situations could have happened," he said, "and it would be interesting if, indeed, both situations are really the same underlying phenomenon."

That is, could the physics of the coldest possible temperature be equivalent to the physics of the hottest possible temperature? Considering that beyond both limits—below one and above the other—space and time start to do those strange, unknown things, Alexander believes it's "a logical conclusion, a logical possibility. Why not?"

Contender #5—Who the heck knows?

As you might have guessed, by this point the physicists had lost me—if not at the very beginning. I was *way* out of my comfort zone.

In the end, perhaps the best answer to my question came from Lee Smolin of the Perimeter Institute for Theoretical Physics in Waterloo, Ontario. "It may be that the most you're going to be able to say is that there's a possibility that there's a highest possible temperature," he told me. "But let me mull it over…."

This feature originally appeared on the site for the NOVA program Absolute Zero.

"Race to Absolute Zero" Video

http://www.youtube.com/watch?v=28F_oPDZHSk&safe=active&noredirect=1

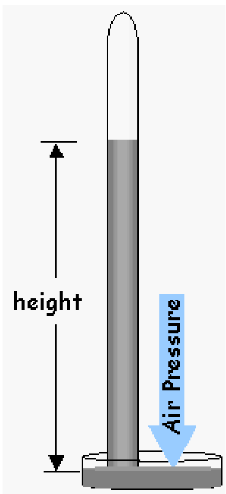
1. What is the closest temperature scientists have gotten to in the race to absolute zero and how did they get there?

2. Choose one scientist and record what they did and when they did it. We are going to add this person to our timeline next class.

3. Describe what matter "looks like" at absolute zero.

The Big Men in Understanding Gases & Pressure

• **1564–1642: Galileo Galilei** He was the first to hypothesize that in a space where we couldn't see anything there must be something there. By inventing a suction water pump he discovered there was a limit to how much water could be "sucked" up by a pump. In order for a pump to lift water, it must decrease in pressure just like we do when "sucking" up liquid thru a straw.



• **1643: Evangelista Torricelli** developed the first barometer. He carried on Galileo's work by determining that the limit to the height Galileo's pump could draw water was due to atmospheric pressure. He invented a closed-end tube filled with mercury that, in turn, was suspended in a shallow dish filled with liquid mercury. The height of the column of mercury in the tube (measured in mmHg) was equal to the atmospheric pressure acting on the mercury in the pan. He determined that the height of mercury supported by atmospheric pressure at sea level is 760 mm or 76 cm.

• **1648: Blaise Pascal** used Torricelli's "barometer" and traveled up and down a mountain in southern France. He discovered that the pressure of the atmosphere decreased as he moved up the mountain. Sometime later the uunit of pressure, the Pascal, was named after him.

1662: Robert Boyle Robert Boyle described what is known as Boyle's law which shows the relationship between the pressure and volume of a gas.

1780: Jacques Charles who expressed the gas law which describes how gases tend to expand when heated.

• **1811: Amadeo Avogadro** suggested, from Gay-Lussac's experiments conducted three years earlier, that the pressure in a container is directly proportional to the number of particles in that container (known as Avogadro's Hypothesis). This can be illustrated by blowing up a balloon, ball, or tire: the more air is added the larger the container becomes due to increased pressure. We learn more about this guy in two units.

Create a timeline describing the understanding of gases. On your timeline place the new information we discovered about gases at that time, and the scientist who discovered it. Include the date (but don't memorize dates) and a mini picture to help you remember.

Gas Laws So Far…..

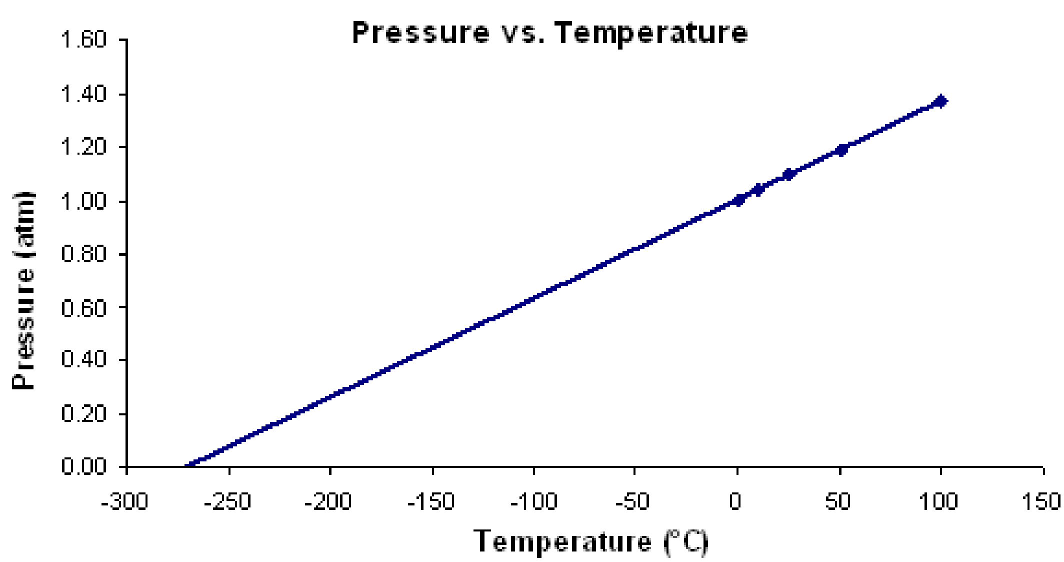
Gas Law Summary

Law	Statement
Boyle's	*P* inversely proportional to *V*
Charles's	*V* directly proportional to T^*

Gay-Lussac's Law

Power Point slides 30 – write what is in blue

Gay-Lussac determined that…



Notice that if we extrapolate the line to 0 pressure, the line crosses the x axis at –273°C , or absolute zero. This is because zero volume would also have zero pressure.

What would the Kinetic Molecular Theory tell us?

Solving Problems using Gay-Lussac's Law

Equation:

Example 1: If a 12.0 L sample of gas is found to have a pressure of 101.3 kPa at 0.0°C, calculate the new pressure at 128°C if the volume is held constant.

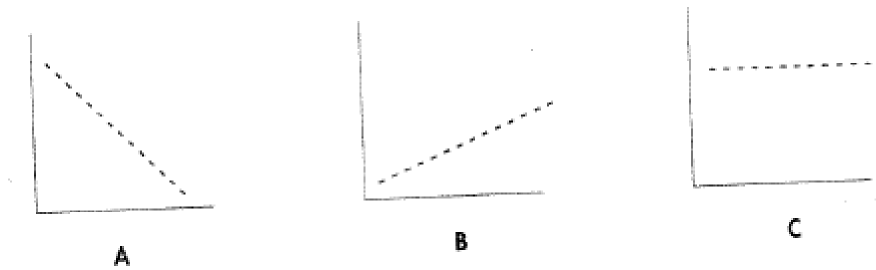
Summary:

-
-
-

http://phet.colorado.edu/en/simulation/gas-properties

Practice Questions:

Match each of the relationships given below with an appropriate graph. Each letter can be used once, more than once, or not at all.



_____ 1. The relationship between gas pressure and its temperature.
_____ 2. The relationship between gas volume and its pressure.
_____ 3. The relationship between gas temperature and volume.

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

As a gas is compressed in a cylinder...
4. the distance between gas molecules ___________________
5. the number of gas molecules___________________
6. its volume___________________
7. its pressure___________________
8. its density___________________
9. its mass___________________

Gay-Lussac's Law Worksheet

Note: STP stands for Standard Temperature and Pressure. Standard temperature is 0 ºC (273 K) and standard pressure is 1 atm (101.3 kPa, 760 mm Hg).

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

$$20.0\ ^{\circ}C = 293\ K,\ 30.0\ ^{\circ}C = 303\ K$$
$$1.00\ atm / 293\ K = P_2 / 303\ K$$
$$P_2 = 1.03\ atm.$$

2. A gas has a pressure of 0.370 atm at 50.0 °C. What is the pressure at standard temperature (273 K)?

50.0 ºC = 323 K
0.370 atm / 323 K = P_2 / 273 K
P_2 = 0.313 atm

3. A gas has a pressure of 699.0 mm Hg at 40.0 °C. What is the temperature at standard pressure (760 mm Hg)?

40.0 ºC = 313 K
699.0 mm Hg / 313 K = 760 mm Hg / T_2
T_2 = 340. K

4. A gas at 750.0 mm Hg is cooled from 323.0 K to 273.15 K. If the volume is kept constant, what is the final pressure?

750.0 mm Hg / 323.0 K = P_2 / 273 K
P_2 = 634.35 mmHg

5. 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 is the final temperature of the gas?

25.0 ºC = 298 K
15.0 atm / 298 K = 16.0 atm / T_2
T_2 = 318 K

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

20.0 ºC = 298 K, 50.0 ºC = 323 K
3.00 atm / 293 K = P_2 / 323 K
P_2 = 3.31 atm

7. A sample of gas at 3.00×10^3 mm Hg inside a steel tank is cooled from 500.0 °C to 0.00 °C. What is the final pressure of the gas in the steel tank?

500.0 ºC = 773 K, 0.00 ºC = 273 K

$$3.00 \times 10^3 \text{ mm Hg} / 773 \text{ K} = P_2 / 273 \text{ K}$$

$$P_2 = 1059 \text{ mm Hg}$$

8. The temperature of a sample of gas in a steel container at 30.0 kPa is increased from -100.0 °C to 1000.0 °C. What is the final pressure inside the tank?

100.0 ºC = 373 K, 1000.0 ºK = 1273 K

$$30.0 \text{ kPa} / 173 \text{ K} = P_2 / 1273 \text{ K}$$

$$P_2 = 221 \text{ kPa}$$

9. Calculate the final pressure inside a scuba tank after it cools from 1000.0 °C to 25.0 °C. The initial pressure in the tank is 130.0 atm.

1000.0 ºC = 1273 K, 25 ºC = 298 K

$$130.0 \text{ atm} / 1273 \text{ K} = P_2 / 298 \text{ K}$$

$$P_2 = 30.4 \text{ atm}$$

Challenge Questions:

1. Tire experts say that driving your car can increase the tire pressure by about 4.5 kPa every 5 minutes over the first 20 minutes of driving. If the air in your tires begins at 15.0°C with a pressure of 182 kPa, what would be the temperature of the air, in degrees Celsius, in your tires after a 20 minute drive?

   T_1= 15°C = 288.15 K

   P_1 = 182 kPa

   T_2 = ?

   P_2 = (4.5 x 4 = 18kPa) = 182 + 18 = 200 kPa

$$\frac{182\ kPa}{288.15} = \frac{200\ kPa}{T_2}$$

$T_2 = 316.66\ K = 43.5\ °C$

2. The maximum pressure of a steel tank is 125 atm. The tank is filled with a cold gas at room temperature, 21°C, and heated to 375°C. What is the maximum cold pressure of the gas such that the maximum pressure is not exceeded?

$P_1 = ?$

$T_1 = 21°C = 294.15\ K$

$P_2 = 125\ atm$

$T_2 = 648.15\ K$

$$\frac{P_1}{294.15\ K} = \frac{125\ atm}{648.15\ K}$$

$P_1 = 56.7\ atm$

Mixed Up Gas Laws

How to know which equation to use:

1. Label your variables within the question.
   a. Match up pressure and volume and temperature with their correct units.

Temperature units –
Pressure Units –
Volume Units –

2. Compare to your gas laws.
   a. Choose your equation

3. Make sure your units work together.
   a. Kelvin
   b. Pressures
4. Give it a go!

Practice:

1. A balloon is 4.2L at the store where it is 25°C and then taken into my car where it is 7°C. What is the new volume?

$V_1 = 4.2L$
$T_1 = 298.15\ K$
$V_2 = ?$
$T_2 = 280.15$

$V_1/T_1 = V_2/T_2$
$4.2/298.15 = V_2/280.15$
$V_2 = 3.9\ L$

2. A volleyball measures 15.6 psi and 3.9L. When we take it into a new gym the pressure now measures 1.2 atm? What is the new volume?

$15.6/14.7 = x/1\ atm$
$X = 1.061\ atm$

$P_1 = 1.061\ atm$
$V_1 = 3.9\ L$
$P_2 = 1.2\ atm$
$V_2 = ?$

$P_1V_1 = P_2V_2$
$(1.061)(3.9) = V_2\ (1.2)$
$V_2 = 3.4\ L$

3. A 6L gas tank records 2.5 atm at 15°C and then is heated and measures 409 kPa. What is the new temperature in degrees Celsius?

$V_1 = 6L$ (don't need)
$P_1 = 2.5\ atm$
$T_1 = 288.15\ K$
$P_2 = 4.034\ atm$ (converted)
$T_2 = ?$

$P_1/T_1 = P_2/T_2$
$(2.5)(288.15) = (4.034)/T_2$
$T_2 = 464.96\ K = 191.81\ °C$

4. A syringe has a volume of 30cc at room temperature (25^{o}C) and a pressure of 101.3 kPa. When it is submerged in water that is 95^{o}C the pressure changes to 825mmHg. What is the new volume? (tough one – if you can do this you will be fine)

V_1 = 30 cc
T_1 = 298.15 K
P_1 = 101.3 kPa
T_2 = 368.15 K
P_2 = 109.001 kPa (converted)
V_2 = ?

$P_1V_1/T_1 = P_2V_2/T_2$
V_2 = 34.4 cc

Combined Gas Laws

Now that we have solved problems using each gas law individually, we can use them together.

Remember the relationships:

- Boyle's Law: _Pressure and Volume___ are ___indirectly___ related.
- Charles' Law: __Volume and Temperature___ are _directly__ related.
- Gay-Lussac's Law: __Pressure and Temperature__ are __directly__ related.

Using Equations ….

Knowing what you do about the three laws see if you can combine them into a **combined gas law**.

Law	Variables	Relationship	Equation
Boyle's			
Charles'			
Gay-Lussac's			
Combined			

			$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ • **Temperature must be in Kelvins and the amount of gase must stay the same**

<u>Here are the steps we want to follow:</u>
1. What do you have?
2. What are you looking for?
3. Convert to Kelvin.
4. Set up the equation you need by rearranging the variables.
5. Watch to keep your units the same.
6. Good luck!!!! Enjoy!!!!

Example 1:
If a gas occupies a volume of 25L at 25^{o}C and 1.25 atm, calculate the volume at 128^{o}C and 0.750 atm.

Example 2:
A gas has a volume of 125 L at 325 kPa and 58^{o}C, calculate the temperature in Celcius to produce a volume of 22.4L at 101.3 kPa.

Combined Gas Laws Problems

1. Carbon dioxide occupies a 2.54 dm^3 container at STP. What will be the volume when the pressure is 150 KPa and 26^{o}C?

2. Oxygen occupies a fixed container of 5.5L at STP. What will happen to the pressure if the temperature rises to 300K?

3. Methane is compressed in a closed 15.8 dm^3 container at 101.3 KPa. If the volume drops to 8.7 dm^3 and the temperature begins at 25°C and then drops to 18°C , what will the pressure of the gas be?

P_1 = 101.3 k Pa
V_1 = 15.8 dm^3
T_1 = 25 ºC = 298 K

P_2 = ?
V_2 = 8.7 dm^3
T_2 = 18 ºC

$$\frac{101.3\text{ kPa} \times 15.8\text{ dm}^3}{298\text{ K}} = \frac{P_2 \times 8.7\text{ dm}^3}{291\text{ K}}$$

$$P_2 = 180\text{ kPa}$$

4. A helium balloon is fully inflated at 2.2 L. When the clerk is filling the balloon, she stops to make sure it is not going to explode and checks the pressure. The pressure is 7.65 mm Hg and the temperature at the store is 24°C. The balloon is 2.0L. She stops when the balloon is fully inflated and the pressure is 775 mm Hg. When the customer takes the balloon outside, it explodes. What was the temperature outdoors?

P_1 = 765 mm Hg
V_1 = 2.0 L
T_1 = 24ºC = 297 K

P_2 = 775 mm Hg
V_2 = 2.2 L
T_2 = ?

$$297\text{ K} \times \frac{2.2\text{ L}}{2.0\text{ L}} \times \frac{775\text{ mm Hg}}{765\text{ mm Hg}} = 330\text{ K}$$

The temperature is at least 330 K. This is pretty hot. It is 60 ºC.

5. Oxygen gas is added to a rebreather at the same rate as it would be in the open air. The volume is 4.8 dm^3 and the temperature is 25 °C. The pressure is1 atm. Once the tank is lowered under the water the pressure increases to 2 atm and the temperature drops to 12°C. Is the tank in danger of exploding if its maximum volume is 5.0 dm^3?

$$4.8\text{ dm}^3 \times \frac{1.0\text{ atm}}{2.0\text{ atm}} \times \frac{285\text{ K}}{298\text{ K}} = 2.3\text{ dm}^3$$

The tank is not in danger of exploding.

7. You are researching the gases that keep lily pads afloat. The cell vacuoles that contain the oxygen that keeps the plant afloat are approximately 17.7 μL. They can expand and contract with changes in temperature and pressure, but that is the optimal volume at 25 ^{o}C and standard pressure. If the volume drops below 14.2 μL the leaves are submerged. You observe the lily pads in the pond are below water level, so the volume is down. If the pressure is 98.2 KPa, what is the temperature?

$$\frac{101.3\text{ kPa} \times 17.7\ \mu\text{L}}{298\text{ K}} = \frac{98.2\text{ kPa} \times 14.2\ \mu\text{L}}{T_2}$$

$$T_2 = 231.8\text{ K} = -46.4\ ^{o}\text{C}$$

This temperature is only likely during the winter when the plant is not growing, so it is not likely to be a problem.

8. You like to skin dive. The best corals are at a depth of approximately 10m. You take a deep breath and dive. Your lung capacity is 2.4L total. The air temperature is 32 ^{o}C and the pressure is 101.3 kPa. At 10 m the temperature is 21 ^{o}C and 141.2 KPa. What is the volume of your lungs?

9. If you breathe 3.0 L of helium at 25 ^{o}C and 101.3 KPa, you will talk funny. You think that would be fun. You breathe all the helium in a container at 15 ^{o}C and 110.6 KPa and you aren't talking funny. Why not?

$$\frac{101.3\text{ kPa} \times 3.0\text{ L}}{298\text{ K}} = \frac{110.6\text{ kPa} \times V_2}{288\text{ K}}$$

$$V_2 = 2.7\text{ L}$$

It is not enough volume of gas to talk funny.

10. A researcher is studying the relationship between volume of burps and stress. Burps due to stress are just swallowed air. Your diaphragm detects pressure as discomfort at 210.0 KPa. (body temperature is 37 ^{o}C). If you feel discomfort and you burp, into the researcher's balloon, the balloon slightly inflates to 6.45 mL at 101.3 KPa and 27 ^{o}C. What volume did the burp occupy in your body?

$$6.45 \text{ mL} \times \frac{101.3 \text{ kPa}}{210.0 \text{ kPa}} \times \frac{310 \text{ K}}{300 \text{ K}} = 3.22 \text{ mL}$$

Gas Activity

In this activity we will explore the effects of temperature on the pressure and volumes of gases.

Materials:
Balloon
Erlenmeyer flask
Water
Hotplate
Ice water bath (large beaker and ice water)
Beaker tongues

Procedure:
1. Pour approximately 30 ml of water into your flask.
2. Put your balloon over the mouth of the flask.
3. Heat your flask on a hot plate. Use tongues when moving your flask on and off because it will be hot!
4. Leave the flask on the hot plate until the balloon is fully upright.

What effect is the increase in temperature having on the liquid in your flask? ___________
____
____

What effect is the increase in temperature having on the gas in your flask? ___________
____
____

Why is the balloon inflating? (Hint-there are 2 reasons) ____________________________
____
____

5. Lift your flask (with tongues) off of the hot plate and put in in the ice water bath.

What happened? ____

Use the Kinetic Molecular Theory to explain why? ______________________
____
____
____

Practical Applications of Gases

Air bags:

Air bags are a supplementary restraint system found in vehicles. An air bag slows a passenger's speed to zero with generally little or no damage. In the air bag's inflation system, sodium azide, NaN_3, reacts with potassium nitrate, KNO_3, to produce a rapid blast of nitrogen gas.

What happens to the volume of the air bag upon collision?
____

Why does it make more sense to have air bags as opposed to say water bags? Think about the Kinetic Molecular Theory. Include at least 3 reasons.
____
____
____
____

If the pressure of the air bag before the collision was 8 atm and the pressure after the collision was 1 atm. The final volume was 25 L, what was the volume of the air bag before the collision?

Propane Tanks (BBQ anyone?)

http://www.energyeducation.tx.gov/fuels/section_3/topics/propane/index.html

Propane appliances: appliances using propane as a fuel source. Propane often serves as a fuel for home heating systems. It is also used for a variety of applications, including cooking, clothes drying, pool/sauna heating, fireplaces, and grilling. The propane refrigerator found in many recreation vehicles is an application of propane in an appliance.

Explain using a Gas law and the Kinetic Molecular Theory why propane is a liquid in the tank and a gas when we burn it. Be specific. ________________________________
____
____
____
____
____
____

Draw your own diagram demonstrating the gas law and the state change.

SCUBA Divers

Self-contained underwater breathing apparatus (scuba): an apparatus that divers use to provide them with air or other breathable gases while they are submerged.

All three gas laws have an effect on underwater divers who must take into consideration all their combined effects when planning a dive. It is very important that the pressure within the gas tank is equal to the ambient pressure (the pressure at whatever depth they are at). This means that at 30 m the ambient pressure is 4 atm, a regulator on the tank will adjust the pressure within the tank to achieve 4.0 atm.

As a diver descends (goes down) does the pressure increase or decrease? ____________________
What is causing the change in pressure? ____
____

The danger for a diver on ascent (coming back up) is that the gases within the body expand forming air bubbles.
Why would the gases expand? ____

If we need to decrease the volume of the gas bubble what do we need? ____________

These air bubbles can be forced into the tissues and bloodstream of a diver with life-threatening consequences. To minimize this problem divers wear a buoyancy-compensating device (BCD). This device holds the compressed air tank in place and controls buoyancy. The device inflates and deflates with air to control the rate that a diver descends and ascends. The addition of too much air can cause a diver to accelerate toward the surface as the air inside the BCD expands with the reduced pressure. If the dissolved gas in the bloodstream cannot escape quickly enough, physical damage may result for the diver. So the device must slowly bring the diver to the surface to give the gas in the diver's blood stream time to slowly adjust to the pressure. Too fast and damage can be done.

Explain why the speed at which a diver comes back up is important; relate it to Boyle's Law.

____
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Our Atmosphere: From Then till Now

1. Listen to our fable

Three equations helped form our atmosphere so that we could live on earth.

Research the following three processes. Include chemical reactions for them and a brief explanation of why they are important. You will be asked to explain one on the test.

Photosynthesis:
Early organisms called cyanobacteria were able to take carbon dioxide and water from the ocean and turn it into sugar and oxygen (O_2) by photosynthesis.

Photodecomposition:
This oxygen gas was released into the atmosphere. More oxygen gas was added to the atmosphere as ultraviolet light broke down water vapor.

Formation of ozone:
Once there was enough oxygen gas it began reacting with itself to from ozone (O_3) which protected the surface from the harmful effects of ultraviolet light.

2. What was the moral of the fable?

"Ok here comes the moral to the story - Earth's atmosphere is still changing. The amounts of nitrogen, oxygen and argon are not changing much. But human activities are changing the amounts of carbon dioxide, sulfur dioxide, nitrogen oxides, methane, ozone at ground level and chlorofluorocarbons (CFCs) that are all increasing. Many of these are greenhouse gases, and contribute to warming the atmosphere. We need the greenhouse effect to keep Earth warm. Without it, life would probably not have evolved. But as the amounts of greenhouse gases in the atmosphere increase, so does the greenhouse effect and the average temperature of Earth.

We don't yet know what all of the effects will be. It's our choice. What will we do?"

The moral is ____
____

3. Greenhouse Effect: Good

Explain why the greenhouse effect is a really good thing. Feel free to use the internet to help you research this. Please draw a diagram to help.

____
____
____
____
____
____

This website has a neat animation that shows how and why CO_2 reflects light and heat energy back. Give it a try. http://phet.colorado.edu/en/simulation/greenhouse

Greenhouse Effect: Bad

Ok so now the heavy stuff…..

Why does the greenhouse effect have such a bad reputation in the media? ____________
____
____

How is the greenhouse effect linked to global warming? __________________________
____
____

____
____

What are human’s doing to cause global warming and climate change? _______________
____
____
____
____

What are some thinkgs we can do to help? ____
____
____
____
____

When you finish this please have me come around and check it with you.

Then please go ahead and start your gases test review.

Gases Review Dictionary

Term	Definition	In your own words (could be another language)	Example or picture or diagram
Pressure			
Atmospheric pressure			

Kelvin			
Boyle's law			
Charles' Law			
Gay-Lussac's Law			
Volume			
Absolute Zero			
Direct relationship			

Indirect relationship			
Manometer			
Kinetic Molecular Theory			
Bose-Einstein Condensate			

Gases Test Review

What you should do to prepare for the test:

1. Review over your notes. Ask questions about anything you do not understand.
2. Write your own definitions for the terms above so that you can explain them on the test.
3. Make sure you have completed all the worksheets and handouts given in class. If not do them as practice now. If so go over some of the questions you had trouble with and do it again. The answer keys are in the binder.
4. Go over your practice quiz. Those are great test questions.
5. Get a good night's sleep and have a good breakfast. You will do great!

You must be able to:

- Describe in point form the four stages of development Earth's atmosphere has undertaken. Include the types of gases and how those gases were produced.

- Know the scientists who have contributed to the Gas Theory and what their contributions were. You do not need to know dates. Write brief sentences beside each of them to show you know their contributions.
    - Galileo
    - Torricelli
    - Pascal
    - Avogadro
    - Boyle
    - Charles
    - Gay-Lussac
- Be familiar with the various units used to measure pressure and be able to convert form one to another.
- Be able to answer questions about pressure from manometers.
- Boyle's Law
    - **Explain** the relationship between pressure and volume.
    - Give an example of Boyle's Law.
    - Do calculations involving Boyle's Law.
    - What is the mathematical equation that represents this law?
- Charles Law
    - **Explain** the relationship between temperature and volume.
    - Explain absolute zero and its importance.
    - Give an example of Charles' Law.
    - Do calculations involving Charles' Law.
    - What is the mathematical equation that represents this law?
- Gay-Lussac's
    - **Explain** the relationship between pressure and temperature.
    - Give an example of Gay-Lussac's Law.
    - Do calculations involving Gay-Lussac's Law.
    - What is the mathematical equation that represents this law?
- Be able to calculate combined gas law problems. You may use the equation or the ratio to calculate your answers.

Practice Gas Law Questions

1. What is the pressure of a contained gas in kPa and psi, if it is at a pressure of 2.5 atm?

2. What is the pressure of a balloon (in millibars) that is attached to a monometer whose mercury levels are equal at an atmospheric temperature of 2 atm?

3. What is the pressure of a flask that is attached to an open ended manometer? The mercury is 22mm higher on the open ended side than on the flask side. The atmospheric temperature is 780 mmHg.

4. We went undefeated in our Friday night volleyball tournament (we beat Crocus and Westwood), but then I went and forgot the balls in the backseat of my car. During our games the balls were at a pressure of 110 kPa, had a volume of 5L, and the temperature in the gym was 28^{o}C. The next day when I pulled the balls out for our warm-up they had gone flat. The volume was the same but the pressure had decreased to 95 kPa. What was the temperature outside?

5. The Trojan basketball team were practicing (before provincials-it is a big deal) when they noticed the balls were bouncing higher than before. The temperature in the gym had gone up 2^{o}C from 21^{o}C to 23^{o}C. What happened to the pressure of the balls if they started at 1 atm? What happened to the volume of the balls if they started at 6.2 L?

6. I filled balloons for my birthday, but when I carried them up the Eiffel Tower (that's were I was celebrating) they burst. How come?

7. In case you were wondering the balloons started at a volume of 3500 mL and burst when the volume reached 4500 mL. The pressure on the ground was 1 atm, what was the pressure at the top of the Eiffel Tower? (I just made these numbers up – I did not really take balloons to the Eiffel Tower.)

8. What happens to the volume of a balloon that begins at a temperature of 25 degrees Celcius, at a pressure of 121 kPa, and a volume of 3.2 L, in the following situations?

    a. The pressure is doubled.

    b. The temperature is doubled from 25oC to 50oC.

    c. The pressure decreases to 0.75 atm.

    d. The pressure becomes 105 kPa and the temperature becomes 320 K.

9. The vapour pressure above a pot of water at 56oC was 2.2 atm and occupied 1L of space. The temperature was then increased to the boiling point of water and a new lid was put on the pot that gave the gas 1.5L of space. What is the new pressure?