611-2300 (30-170) Gas Law Apparatus

Introduction:

This apparatus is designed to allow the student to verify the Ideal Gas Law, **PV = nRT**

It consists of a syringe held between two blocks.

The experiment has two parts.

1) Verifying change in volume of a gas with changes in pressure at constant temperature (Boyles' Law)
2) Verify change of volume of a gas with changes in temperature at constant pressure (Charles' Law)

By combining the two relationships, we arrive at the ideal gas law.

Additional Materials Needed:

- **Uniform weights**
- **Vernier calipers (preferred) or ruler**
- **Large beaker**
- **Thermometer**

Warranty:

We replace all missing or defective parts free of charge. All products guaranteed free from defect for 90 days.This guarantee does not include accident, misuse, or normal wear and tear.

Assembly Procedures

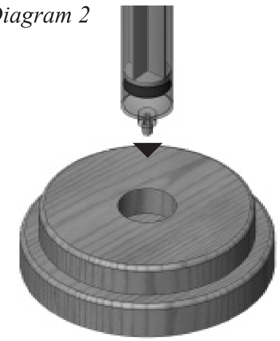
Syringe Assembly

1. With the plunger extended, attach the cap to the end of the syringe, capturing a volume of air to be compressed (at least 20 cc). *See Diagram 1.*

Diagram 1

2. Insert tip of syringe with cap attached into stepped round block. See *Diagram 2.*

Diagram 2



3. Insert plunger head into larger slot on the cap block. See *Diagram 3 and 4.*

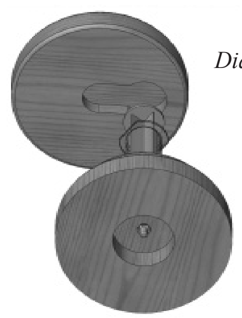


Diagram 3

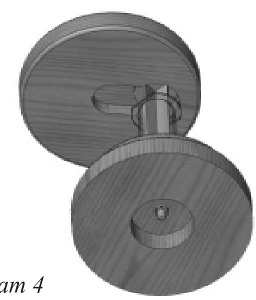


Diagram 4

4. Slide plunger head over to the center of the cap block. The completed assembly is illustrated in *Diagram 5.*

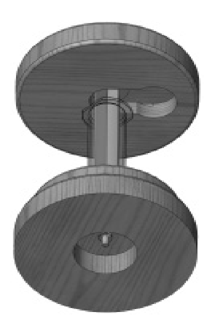


Diagram 5

Verification of Boyles' Law:

Robert Boyle in 1662 first studied the relationship between pressure and volume of a gas at constant temperature. He found that the pressure **P** of the gas was inversely proportional to its volume **V**. In other words, if you double the pressure on a given amount of gas, its volume will halve.

1. Draw 35 cc of air into the syringe and place the cap on the end.
2. Place the syringe into the two blocks as shown.
3. Place uniform weights (identical books work well) on the top block. After placing each weight, record the volume of air in the syringe and the number of weights added. Tapping the syringe lightly will help reduce errors due to friction.
4. After all the weights have been added and the data recorded, remove them one by one, recording the volume and number of weights as before.
5. Weigh each of the weights and record the average weight.
6. Disassemble the apparatus and remove the plunger from the syringe.Using vernier calipers (preferred) or a ruler, measure the inside diameter of the syringe barrel. Calculate the area, **A**, using the formula:
   $$\mathbf{A = \pi r^2}$$
   where **r** is 1/2 of the diameter.
7. Calculate the pressure, **P**, of each data point by taking the average of the two volumes for each run, using the relationship:
   $$\mathbf{P = F/A}$$
   where **F** is the gravitational force exerted by the weight and **A** is the area of the syringe barrel.
8. Plot the volume, **V**, versus **1/P** for each data point. Record the temperature T.

Verification of Charles' Law:

1. Heat a beaker full of water to 90°. Beaker must be large enough to immerse completely the syringe barrel.
2. Draw 30 cc of air into the syringe and cap the end.
3. Place the syringe and a thermometer into the water, using the provided block as a support. The thermometer fits in the smaller hole.
4. Allow 3 to 5 minutes for equilibration.
5. Record the volume in the syringe as well as the temperature.

To get an accurate volume measurement, push on the piston and then release it. Record the volume. Then pull back the piston and release it. Record this second volume and average the two volumes.

6. Allow the water to cool and take successive volume and temperature readings every 10° C or so. To speed cooling, you can add ice to the beaker when the temperature is below 35° C or so.
7. Plot the volume, **V**, versus the temperature **T.** You should get a straight line.

Determining gas constant R:

We know that **P** is inversely proportional to **V** or:

$$\mathbf{V \propto 1/P}$$

and that volume is directly proportional to **T:** $\mathbf{V \propto T}$

We can combine both relationships and write: $\mathbf{V \propto T/P}$

Introducing a constant **K**, we now have: **PV = KT**

But the constant **K** is related to the amount of gas present times the gas constant **R**. We express the amount of gas present in terms of the number of moles present, **n**. Therefore:

$$\mathbf{PV = nRT}$$ where:

P is pressure in atmospheres (atm)
V is the volume in liters
T is the absolute temperature in degrees kelvin
n is the number of moles of gas present.

Rearranging,

$$\mathbf{R = \frac{PV}{nT}}$$

R has the units of l-atm/moles °K.

We now calculate **R** by plugging in the quantities we measured for **P, V** and **T.** We get **n**, the number of moles present, by recognizing that air is made up of 79% N_2 and 21% O_2 by weight. (We will ignore other constituents of air for now.)

The density of air is about 1.20 g/l at 20° C. So 35 ml of air will weigh 1.20 g/l x .035 l = 0.042 g. We calculate "molecular weight" of air by taking the weighted average of the constituent parts of air, as follows: 0.79 (28.0) + 0.21 (32) = 28.8 g/"mole"

where:

28.0 is the molecular weight of nitrogen, N_2 and
32.0 is the molecular weight of oxygen, O_2

Therefore 35 ml of air is about:

$$\frac{.042}{28.8} = 0.00146 \text{ "mole" of air}$$

Try using the data you collected for the Boyles Law experiment to calculate the gas constant, **R.** Repeat using Charles Law data.
The accepted value for **R** is:
0.082 l-atm/mole - ° K
This little demonstration can be surprisingly accurate. Check out the sample run we performed below.

Sample Run

Syringe diameter = .937 inches
Area of syringe is then **A** **= π r²**
= 3.14 (.937/2)^2
= 0.689 sq in
Since Pressure = Force/ Area, and Area times Pressure = Force,
14.7 lb/in^2 x 0.689 in^2 = 10.1 lb/ atm
We then divide the force in lb by 10.1 to get pressure in atm.
To this we add 1.0 atm to arrive at total pressure.

We calculate the number of "moles" of air as:
35 cc = .035 l
Density of air is 1.20 g/ℓ. There are 1.20 g/ℓ x .035 ℓ = 0.042 g of air.
This equals:
0.042 g/28.8 g/mole = 0.00146 "mole" of air.

The experiment was done at 25° C which is
25 + 273° = 298° K.
So the product of **n** and **T** is:
.00146 x 298 = 0.43458
We convert syringe volumes to liters by dividing by 1000.

Vol (ℓ)	**Force (lb)**	**Pressure (atm)**	**P•V**	**T•n**	**R**
.035	0	1.00	.0350	.434	0.0806
.024		5	1.49	.0357	.434
0.0822					
.018	10	1.98	.0356	.434	0.0820
.015	15	2.48	.0372	.434	0.0857
.013	20	2.97	.0386	.424	0.0889
.010	25	3.46	.0346	.434	0.0797

Related Products:

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- **673-0110 Mole Box -** A white cardboard cube, each side 11.2 square inches in size, imprinted with useful information large enough for an entire class to see. Helps to visualize what a "mole" of air "looks" like.
- **611-2025 Ten Density Cubes -** This very popular set contains copper, brass, steel, aluminum, acrylic, oak, nylon, pine, and poplar cubes
- **611-2105 Reverse Density Rod -** Our aluminum rod (611-2100) sinks in hot water, floats in cold. This new-to-industry plastic rod does just the opposite - floats in hot water, sinks in cold. Made of high density polyethylene, 1/2" in diameter. With instructions.
- **611-2020 Four Density Cubes -** One-inch cubes of aluminum, steel, copper and brass include instructions. Use for a variety of labs in density, specific gravity and specific heat.
- **611-2000 Density Set -** Ask your class to identify each of twelve different samples by finding their density. Each cylinder varies ins ize and density since it is constructed of a different material. All have the same diameter but vary in length.

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