1 Gas laws

EXPERIMENT

Gas Laws

Goal

The goal of this laboratory experiment is to verify two of the laws that rule the behaviour of ideal gases, Charle's law and Boyle's law.

Materials

- ☐ Magnesium ribbon
- ☐ Bunsen burner
- ☐ metallic Zn and Cu
- ☐ a 1M $CuSO_{4\,(aq)}$ solution
- ☐ an 1M $HCl_{(aq)}$ solution
- ☐ a series of 0.1M $CaCl_{2\,(aq)}$,$Na_3PO_{4\,(aq)}$,$FeCl_{3\,(aq)}$, and $KSCN_{(aq)}$ solutions
- ☐ $Na_2CO_{3\,(s)}$
- ☐ a 3% $H_2O_{2\,(aq)}$ and 0.1M $KCl_{(aq)}$ solution

Background

Gases and its properties

Gases contain atomic or molecular particles. They have very different properties than liquids or solids. The particles of a gas are spread and far away from each other. Liquids, on the other hand are made of loose particles that interact by means of weak forces. Solids on the other hand are packed materials and its particles, atoms or molecules, are closer together. This section covers the different properties of gases.

Some of the elements in the periodic table are molecular gases, resulting of the combination of two atoms of the same element. For example, molecular oxygen (O_2) is gas. Similarly, molecular nitrogen (N_2), molecular hydrogen (H_2), molecular chlorine (Cl_2), or molecular fluorine (F_2) are all diatomic gases–they contain two atoms of the same element. Other gases result of the combination of two different non-metals. Examples are carbon monoxide (CO) or dioxide (CO_2), and nitrogen monoxide (NO) or dioxide (NO_2). The nobel gases (Ne, He, Ar) also exist in gas state.

Gases has different properties compared to solids of liquids:

- × Gases assume the volume and shape of its container. As they expand, they have no shape different than their container's shape.
- × Gases are compressible: they can be compressed, reducing its volume. Differently, liquids and solids are incompressible.
- × The density of gases is small, compared to the one for solids and liquids.

The volume of a gas (V) is the amount of space it occupies, and gases fully occupy the volume of its container. Liters (L) and milliliters (mL) are units of volume. Liter is a cubic unit and one litter equals to a cubic decimeter (dm^3).

The temperature (T) of a gas is related to the speed (the average velocity) of its particles. The higher the temperature the higher the particles' speed. Although there are different units of temperature such as Kelvin (K) or celsius (C°), in this chapter many formulas require the use of Kelvin temperature (T_K), that related to celsius (T_C) by the formula

$$\boxed{T_K = T_C + 273}$$

The amount of a gas (n) refers to the quantity of gas particles. The larger the amount of gas, the larger the amount of gas particles. The amount of gas is measured in moles or grams. In general pressure is defined as force divided by surface. In the international system the unit of force is the Newton (N) and the unit of area (A) is the m^2. One newton is $1\text{kg}\cdot\text{m/s}^2$.

$$\boxed{P = \frac{F}{A}}$$

The particles of a gas are constantly moving. On its movement, they frequently hit the walls of its container, like the raindrops hitting the ceiling when it rains. When they hit the walls they exert a pressure, and pressure is defined as the force acting on certain area. The larger pressure the stronger the collisions with the walls and the higher the frequency of collision–the stronger the force applied on the walls. Imagine you are driving a motorcycle. While you drive you can feel the collision of the air's particle with your face. The faster you go the higher pressure. The value of air's pressure is measured with a barometer and depends on your location on the earth–in particular your altitude–as well as the weather. If you are at the sea level the atmospheric pressure is one unit of pressure (one atm), due to the air that you have on top of you. If you climb a mountain, the pressure decreases, as there is less air on top of you. The higher you are with respect to the sea level, the lower the air pressure. The weather also affects pressure, and in hot days the pressure of air is higher, whereas on a cold day pressure is lower.
Units of pressure are: bars, atmospheres (atm), torr, pascals (Pa) or millimeters of mercury (mmHg). In order to covert pressure units, you can use the following conversion factors:

$$\boxed{\frac{1\text{ atm}}{1.01325\text{ bar}}} \quad \boxed{\frac{1\text{ atm}}{760\text{ mmHg}}} \quad \boxed{\frac{1\text{ torr}}{1\text{ mmHg}}} \quad \boxed{\frac{1\text{ atm}}{101325\text{ Pa}}}$$

one millimeters of mercury (mmHg) is the same as 1 torr. As a note, the name torr knowledges the person who invented the barometer: Torricelli, an Italian physicist.

Two different devices are used to measure pressure, barometers and manometers. Barometer are used to measure specifically the atmospheric pressure and historically, they consist on a glass tube filled with mercury (Hg), inverted on a plate containing more mercury. At the sea level, the height of the mercury columns should be close to 760mmHg. Manometers, on the other hand, are used to measure the pressure of any gas. Manometers consist of a u-shaped tube filled with mercury. There are two types of manometers: open-tube and closed-tube manometers. The pressure exerted by a gas changes level of mercury on both sides of the tube and the difference in height measured as the right minus the left side ($\Delta h = h_{right} - h_{left}$) is related to the gas pressure. For closed-tube manometers–normally used to measure pressured below the atmospheric pressure–when the gas pressure increases the left column of the barometer is reduced and the right column increases. The difference between both columns is related to the gas pressure by means of:

$$\boxed{\text{P}^{closed} = \text{hdg}} \qquad \boxed{\text{P}^{open} = \text{hdg} + \text{P}_{atm}} \tag{1}$$

where:

P is the pressure of the gas in Pa

Δh is the heigh difference in m, measured as $h_{right} - h_{left}$

d is the density of mercury 13593 kg/m^3

g is gravity, 9.8 m/s^2

P_{atm} is the atmospheric pressure close to 101325 Pa

For the open-tube manometer, normally used to measure pressured above the atmospheric pressure, we need to take into account the atmospheric pressure to the gas pressure. For this type of manometers, if the left column is below the right column ($\Delta h = h_{right} - h_{left} < 0$), this means that the pressure of in the gas is below the atmospheric pressure.

Ideal gas law

Ideal gases are gases made of particles without a size (very tinny) that do not interact with each other. The temperature, pressure, volume and number of moles of a gas are not independent. They are related by the ideal gas law. In this section we will introduce this law in two different forms: in terms of volume and in terms of density.

The ideal gas law says:

$$\boxed{PV = nRT} \quad \text{Ideal Gas Law}$$

where:

P is the pressure of the gas in atm

V is the volume of the gas in L

n is the number of moles of the gas

T is the temperature of the gas in K

R is the constant of the gas $0.082\,\frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}}$

Imagine for example that you inflate a balloon with you mouth, introducing air particles into the balloon. While the number of particles inside the balloon grows, its volume will grow too. More particles will collide with the walls of the balloon and hence, the pressure inside the balloon will also increase.

The ideal gas law in terms of density is:

$$\boxed{P \cdot MW = DRT} \quad \text{Ideal Gas Law in terms of D}$$

where:

P is the pressure of the gas in atm

MW is the molecular weight (or atomic weight, AW) of the gas in g/mol

D is the density in $g \cdot L^{-1}$

T is the temperature of the gas in K

R is the constant of the gas $0.082\,\frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}}$

We use this formula when we are questioned about the molar mass or density of the gas. STP conditions refer to standard temperature (273K) and pressure (1 atm) conditions. Working at STP conditions means pressure will be fixed at 1 atm and temperature at 273K.

$$\boxed{\text{1 atm and 273K}} \quad \text{STP Conditions}$$

Change of gas properties

The previous section addressed the properties of an ideal gas. However, as all properties of a gas are related, if we modify one the others will change too. This section covers situations in which one of the gas properties changes (e.g. V changes) and you need to predict the change of another gas property (e.g. P). For example, imagine you compress a balloon with your hand. The temperature and number of moles of the gas inside the balloon are constant, as the balloon is closed and in contact with the atmosphere. Differently, the pressure and volume will change. In particular, the volume will decrease and the pressure will increase. This means that the gas particles will hit the balloon harder and with more frequency. In order to solve problems in which two of the gas variables are kept fixed and the other two are fixed, one needs to apply the ideal gas law at the initial and final state to then divide both formulas. Imagine the situation in which you have a 1L hot air balloon with 1 moles of a gas and you add gas to a total of 5 moles. You want to calculate the final volume after you inflate the volume, knowing the temperature and pressure are kept constant. The initial state corresponds to 1L and 1 moles of gas and the final estate corresponds to an unknown volume and 5 moles. Using the ideal gas formula twice you have:

$$\left.\begin{array}{r} PV_1 = n_1RT \\ \\ PV_2 = n_2RT \end{array}\right\} \quad \frac{PV_1}{PV_2} = \frac{n_1RT}{n_2RT} \tag{2}$$

as some of the variables of the cancel out:

$$\frac{\cancel{P}V_1}{\cancel{P}V_2} = \frac{n_1\cancel{RT}}{n_2\cancel{RT}} \tag{3}$$

and you end up with Avogadros' law. If you plug the numbers into the formula:

$$\frac{1L}{V_2} = \frac{1\text{ mol}}{5\text{ mol}} \tag{4}$$

and you get a final volume of 5L. If temperature and the number of moles of a gas are kept constant the product of pressure and volume will remain constant too. This is the case of the balloon-pressing example. We call this Boyle's Law:

$$\boxed{\frac{P}{V} = c \quad or \quad P_1 \cdot V_1 = P_2 \cdot V_2} \quad \text{Boyle's law}$$

where:

P_1, V_1 are the initial pressure and volume

P_2, V_2 are the final pressure and volume

c is a constant

Imagine a hot air balloon. Air comes in and out of the balloon as the balloon is not closed balloon. Hence the pressure inside the balloon is just the atmospheric pressure. Also as the balloon is in contact with the air, its temperature will be constant, resulting from the thermal equilibrium between the inside of the balloon and the atmoshphere. If you inflate the balloon with hot air, the volume of the balloon and the number of moles are related by Avogadro's law:

$$\boxed{\frac{V}{n} = c \quad or \quad \frac{V_1}{n_1} = \frac{V_2}{n_2}} \quad \text{Avogadro's law}$$

where:

V_1, n_1 are the initial volume and number of moles

V_2, n_2 are the final volume and number of moles

c is a constant

The questions is now, if we increase the pressure at fixed number of moles and pressure, how do we know if the volume will increase or perhaps decrease? Similarly, if for example the number of gas moles increase at fixed pressure and volume, will the temperature of the gas increase or perhaps decrease. We can answer these questions by means of the ideal gas law. If the variables that we need to relate are in the same side of the equation (e.g. P and V) then if one of the variables increase the other will decrease. Differently, If the gas variables to relate are located in opposite sides of the gas law (e.g. P and T) then both will change in the same direction. For example, let us consider the changes of P and V (at fixed n and T). As they are in the same side of the ideal gas law ($PV = nRT$), if P increases V will decrease. Differently, for the change of P and T (at fixed V and n), as both variables are in opposite sides of the ideal gas law ($PV = nRT$), if P increases, T will increase as well.

Procedure

Charle's Law: warming up the gas

- [ ] *Step 1:* – Set a 400mL beaker on a hot plate and use an iron ring to secure it.
- [ ] *Step 2:* – Obtain a 125mL Erlenmeyer flask and make sure it is dry inside.
- [ ] *Step 3:* – In the neck of the flask, place a rubber stopper containing tubing and a pinch clamp.
- [ ] *Step 4:* – Leave the pinch clamp open.
- [ ] *Step 5:* – Use a clamp to secure the flask to a ring stand.
- [ ] *Step 6:* – Set a 400mL beaker on a hot plate and use an iron ring to secure it. Place the flask inside the beaker without touching the bottom of the beaker.
- [ ] *Step 7:* – Add water to the beaker until it reaches the neck of the flask but without reaching the rim of the beaker.
- [ ] *Step 8:* – Begin boiling the water in the beaker, and boil gently for 10 min so that the temperature of water equilibrates with the temperature of the air inside the flask.
- [ ] *Step 9:* – Measure and record in the Results section the temperature of the boiling water.

Charle's Law: cooling down the gas

☐ *Step 1:* – Remove the flask from the hot plate and close the pinch clamp. Under any circumstance close the pinch while the flask is submerged in hot water.

☐ *Step 2:* – Undo the clamp that secures the flask to the ring stand and use it as handle to carefully lift the flask out of the boiling water.

☐ *Step 3:* – Quickly place the flask with the flask-handle inside a large container with cold water while keeping the stopper-end of the flask pointing downward.

☐ *Step 4:* – Keeping the stopper-end of the flask downward, open the pinch-clamp while keeping the flask submerged.

☐ *Step 5:* – At this point, water will start entering the flask, decreasing the volume of air inside the flask. Keep the flask submerged for 10 minutes. At this point the temperature of the cool water-bath will be the same as the temperature inside the flask.

☐ *Step 6:* – Keeping the flask upside down rise or lower the flask until the water level inside the flask is equal to the water level of the large water container. This process equalizes the pressure inside the flask with the atmospheric pressure. At this point, close the pinch clamp. Now you can remove the flask from the large water container.

☐ *Step 7:* – Measure and record in the Results section the temperature of the large water container.

☐ *Step 8:* – Measure and record in the Results section the amount of liquid that entered the flask by pouring the water into a measuring cylinder.

☐ *Step 9:* – Dry very well the flask, and repeat the procedure described above two more times.

Charle's Law: flask volume

☐ *Step 1:* – Now you are going to measure and record in the Results section the total volume of the flask (mind this is not the volume indicated in the flask).

☐ *Step 2:* – Fill completely the flask with water until its rim, and then pouring this water into a large-enough measuring cylinder.

☐ *Step 3:* – Plot the values of the gas volume (vertical axis) and temperature (horizontal axis). The results should give a straight line. Draw a smooth line through the data points.

Boyle's Law

☐ *Step 1:* – This miniexperiment provides values of pressure and volume of a gas at fixed temperature. You will have to process these data.

☐ *Step 2:* – Multiply the pressure and volume values and report the resulting values in the results section.

☐ *Step 3:* – Plot the values of pressure (vertical axis) and volume (horizontal axis). The results should give a slight curve (not a straight line). Draw a smooth line through the data points.

Calculations

1 This is the volume of the flask obtained by filling the flask completely with water and then displacing the liquid into a cylinder.

2 These are the temperature T in Kelvin of the hot and cold baths.

(3) These are the volume of water displaced into the flask, V_{liquid}.

(4) These values are the volume of gas in the flask, V_{gas}:

$$V_{gas} = (1) - (3)$$

(5) Divide the gas volume by its temperature:

$$\frac{V_{gas}}{T} = \frac{(4)}{(2)}$$

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STUDENT INFO

Name: ____________ Date: ________

Pre-lab Questions

Gas Laws

1. A 3 grams sample of Ar at 40°C is placed in a 3L container. Calculate the pressure inside the container.

2. What is the molar mass of a gas if a 3.16 g sample at 0.75 atm and 45°C occupies a volume of 2L.

3. Dinitrogen oxide, used in dentistry, is an anesthetic also called laughing gas. What is the pressure in atm of 0.35 moles of N_2O at 22°C in a 5L container?

4. In a storage area where the temperature has reached 300K, the pressure of oxygen gas in a 15 L steel cylinder is 1 atm. Calculate the volume if the pressure is reduced to 0.5 atm.

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Results
EXPERIMENT

Gas Laws

Charle's Law

1 Flask volume (L), V_f =

		Hot Bath	Cool Bath		
			Replicate 1	Replicate 2	Replicate 2
2	Temperature (K)				
3	Water volume inside the flask, $V_{(liquid)}$ (L)				
4	Gas volume, $V_{(gas)}$ (L)				
5	$\frac{V_{(gas)}}{T}$ (L/K)				

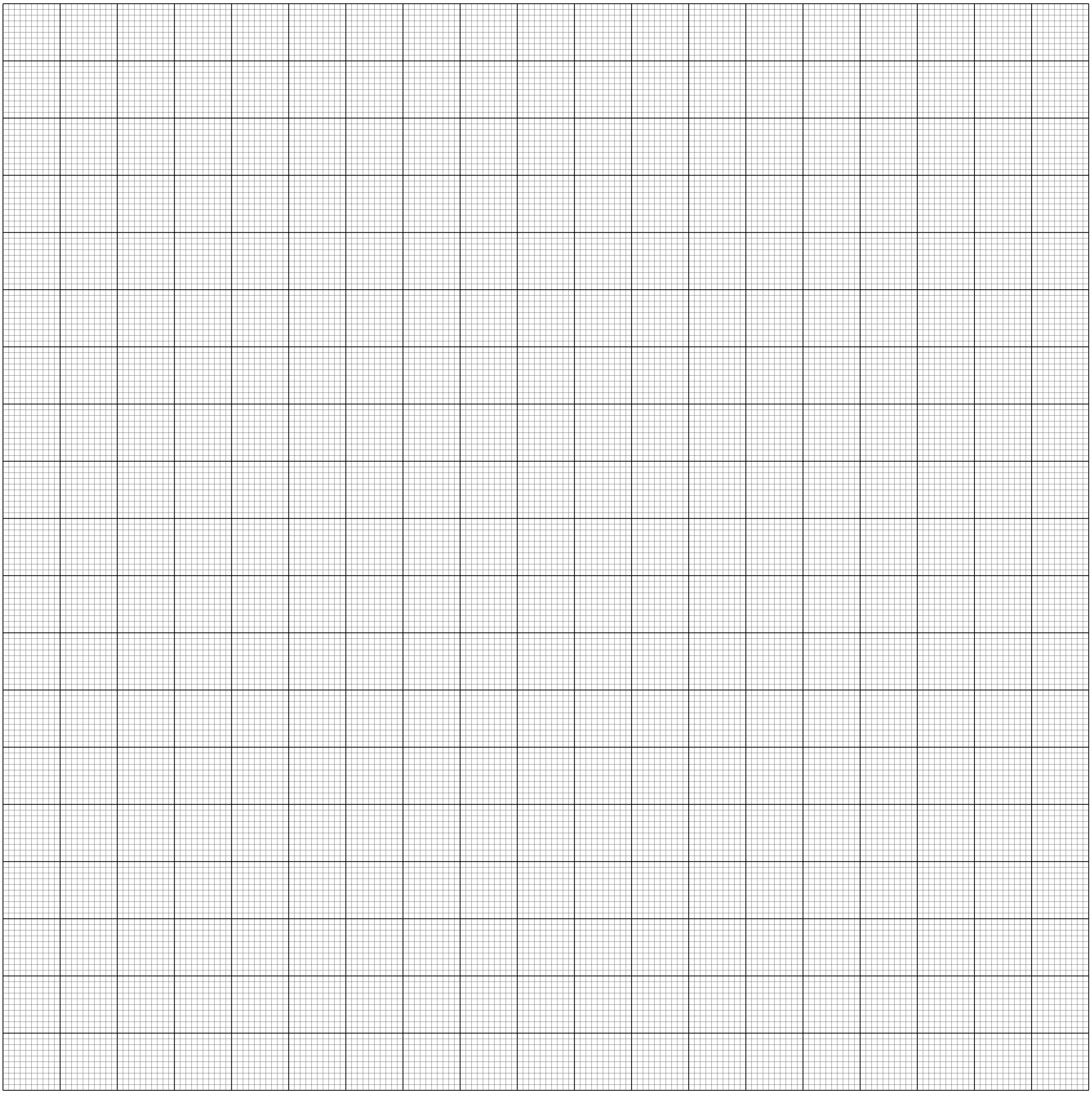

Figure 1: Charle's Law

Boyle's Law

Pressure, P (atm)	Volume, V (L)	P×V (atm· L)
0.5	44.772	
0.6	37.31	
0.7	31.98	
0.8	27.9825	
0.9	24.8733	

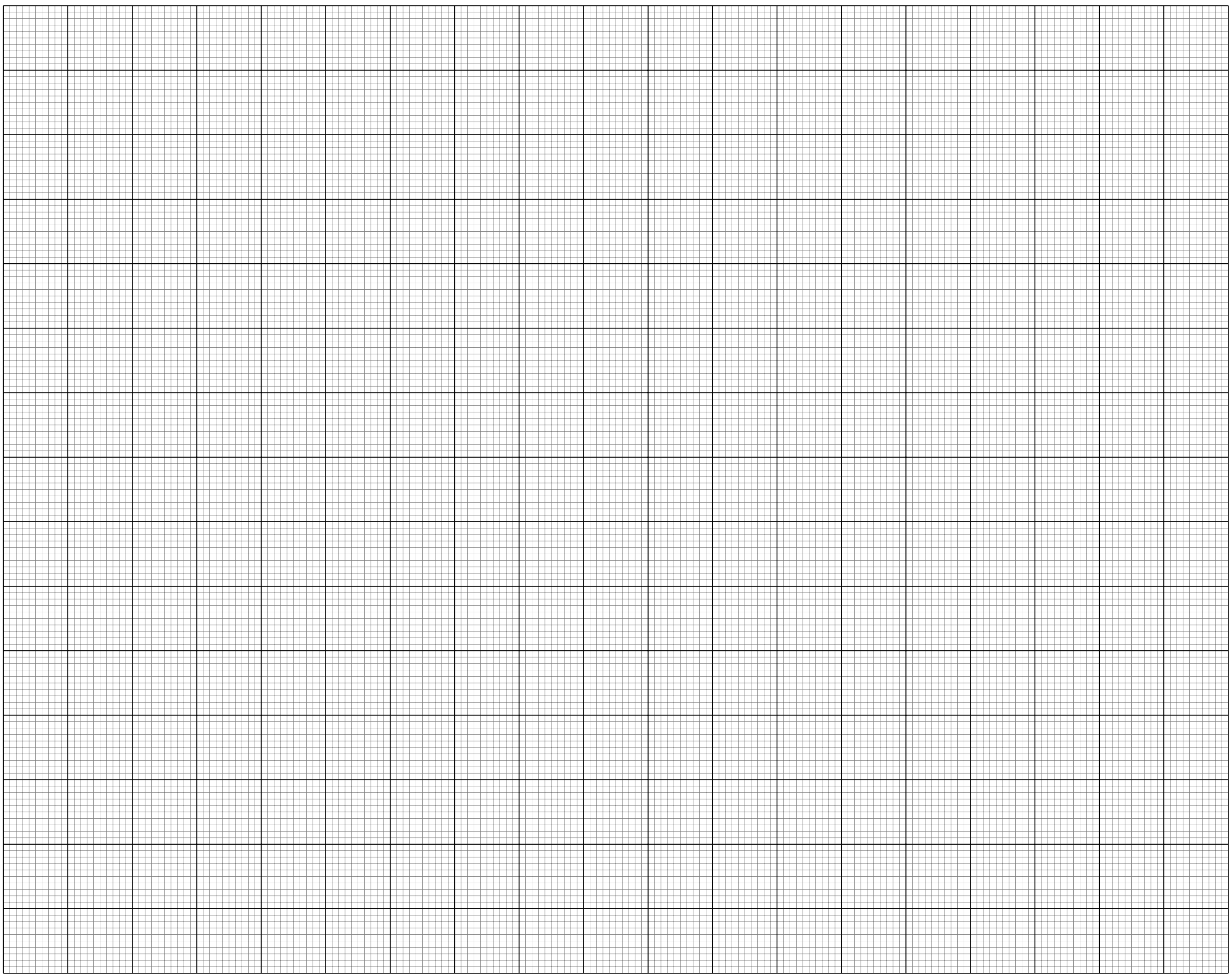

Figure 2: Boyle's Law

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Post-lab Questions

Gas Laws

1. A 3 grams sample of Ar at 40°C is placed in a 3L container. Calculate the pressure inside the container.

2. What is the molar mass of a gas if a 3.16 g sample at 0.75 atm and 45°C occupies a volume of 2L.