Interpreting
UNIVERSAL GAS LAWS

Topic 3: Kinetic theory of gasses

Subtopic(s) 3.07

CRFS question in the 2023 October exam

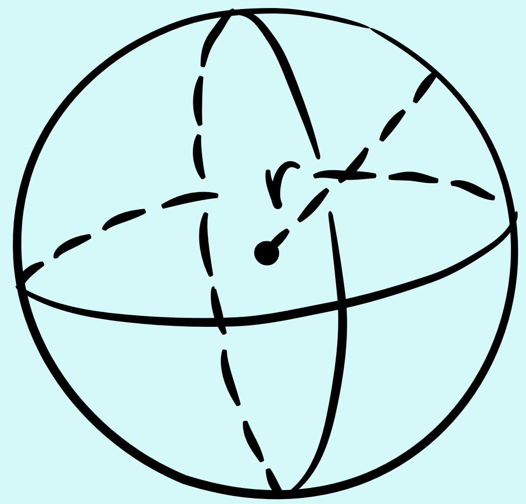
How much does 1L of air weigh?

Universal Gas Laws

Universal gas laws and conversion factors are fundamental concepts in chemistry and physics, particularly relevant in fields like respiratory medicine and engineering where precise measurements and conversions are critical.

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$y = mx + b$$



$$V = \frac{4}{3}\pi r^3$$

Universal Gas Laws

Boyle's Law: States that for a given amount of gas at constant temperature, the pressure and volume are inversely related.

$$P \propto \frac{1}{V} \quad PV = k$$

Charles's Law: States that at constant pressure, the volume of a gas is directly proportional to its temperature (in Kelvin)

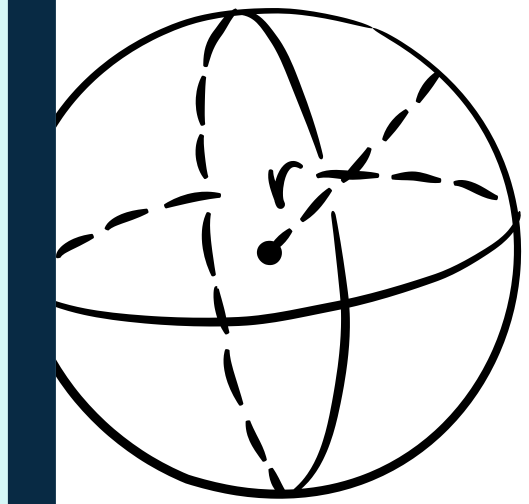
$$V \propto T \quad \frac{V}{T} = k$$

Gay-Lussac's Law : States that at constant volume and amount of gas, the pressure of a gas is directly proportional to its temperature (in Kelvin)

$$P \propto T \quad \frac{P}{T} = k$$

$$\frac{-b \pm \sqrt{b^2 - 4a}}{2a}$$

$$y = mx + b$$



$$V = \frac{4}{3}\pi r^3$$

Universal Gas Laws

Avogradro's Law: states that, at constant temperature and pressure, the volume of a gas is directly proportional to the number of moles of the gas

$$V \propto n \quad \frac{V}{n} = k$$

k as a constant: In each gas law, k serves as a constant that is specific to the conditions being described (such as temperature, pressure, and volume). This constant helps to quantify the relationship between the variables of the gas law under specific conditions

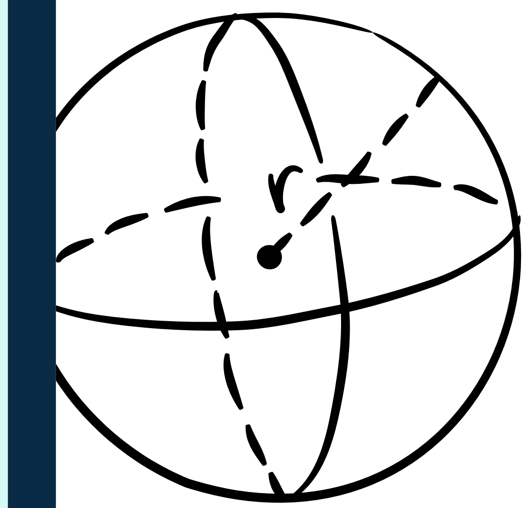
The Ideal Gas Law is applicable to many real gases under a range of conditions, particularly when:

- The gas is at relatively low pressure.
- The gas is at a high temperature.

In these conditions, gases like Nitrogen (N_2), Oxygen (O_2), Hydrogen (H_2), and Carbon Dioxide (CO_2) often approximate ideal gas behavior closely.

$$\frac{-b \pm \sqrt{b^2 - 4a}}{2a}$$

$$y = mx + b$$



$$V = \frac{4}{3}\pi r^3$$

Universal Gas Laws

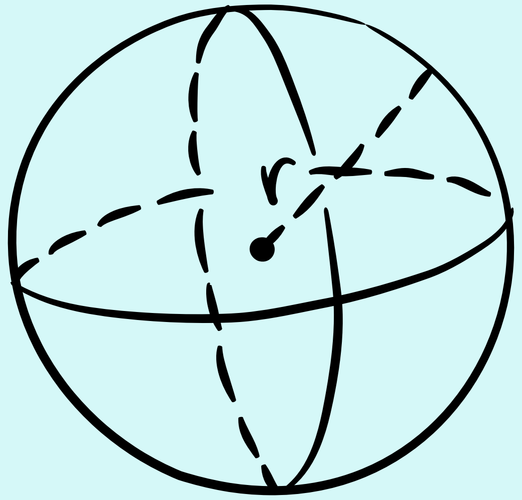
1. Ideal Gas Law

Equation: $$PV = nRT$$

- P = Pressure of the gas
- V = Volume of the gas
- n = Number of moles of the gas
- R = Universal gas constant (8.314 J/(mol·K) or 0.0821 L·atm/(mol·K))
- T = Temperature of the gas

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$y = mx + b$$



$$V = \frac{4}{3}\pi r^3$$

EXPRESSIONS OF UNITS:

Pressure

- Atmospheres (atm)
- Millimeters of Mercury (mmHg)
- Pascals (Pa)
- Kilopascals (kPa)

Volume

- Liters (L)
- Milliliters (mL)

Amount of gas

- Number of moles (n)

Temperature

- Kelvin (K)
- Celcius (°C)

$$PV = nRT$$

$$R = \frac{0.0821 \times atm \times L}{mol \times K}$$

Where is the 0.0821 from?

Avogradro's law:

- At standard temperature and pressure (STP) of 1 atm and 0°C... 1 mol of any gas takes up 22.4L

$$PV = nRT$$

$$R = \frac{1atm \times 22.4L}{1mol \times 273K}$$

$$R = 0.082051$$

Conversion of Pressure Units

kPa

Kilopascal is a thousand 'pascals' (unit of pressure). kPa aligns with the international systems of Units (SI)

1 atm = 101.3kPa

mmHg

This unit comes from an old way of measuring pressure using a column of mercury

1 atm = 760mmHg

EXAMPLES WITH OTHER UNITS OF PRESSURE:

kPa

$$R = \frac{101.3kPa \times 22.4L}{1mol \times 273K}$$

$$R = \frac{8.31kPa \times L}{mol \times K}$$

mmHg

$$R = \frac{760mmHg \times 22.4L}{1mol \times 273K}$$

$$R = \frac{62.4mmHg \times L}{mol \times K}$$

Why do we use Kelvin?

Absolute zero reference

- Kelvin is an absolute temperature scale. It starts at absolute zero (0 K), which is the theoretical point where atomic motion stops and complete absence of thermal energy.

Direct proportionality

- The size of the degree is the same as in Celsius, but it avoids negative temperatures. having a scale that starts at zero and only uses positive values simplifies the math

Scientific consistency

- Gases (like the ideal gas law) or thermodynamics, are based on absolute temperature. Using Kelvin ensures consistency and avoids conversion errors.

Standardization

- Kelvin is the standard unit of temperature in the International System of Units (SI). It's important to use standardized units to avoid confusion and ensure accuracy

CONVERTION OF TERMPERATURE & VOLUME

Temperature

$$0^\circ C = 273K$$

$$25^\circ C = 298K$$

Volume

$$1000mL = 1L$$

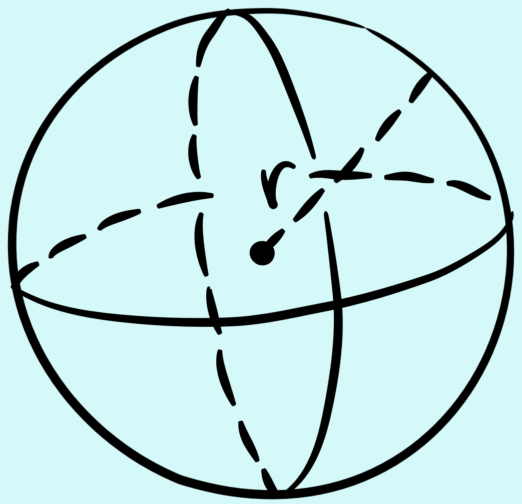
$$2500mL = 2.5L$$

Question 1:

2.3 moles of helium gas are at a pressure of 1.70 atm and the temperature is 41°C. What is the volume of gas?

$$\frac{-b \pm \sqrt{b^2 - 4a}}{2a}$$

$$y = mx + b$$



$$V = \frac{4}{3}\pi r^3$$

ANSWER:

2.3 moles of helium gas are at a pressure of 1.70 atm and the temperature is 41°C. What is the volume of gas?

1. Ideal Gas Law

Equation: $PV = nRT$

- P = 1.70atm
- V = ?
- n= 2.3mol
- R= Universal gas constant (0.0821 L·atm/(mol·K))
- T = 41°C ----> (+273) **314K**

2. Rearrange equation

$$V = \frac{nRT}{P}$$

$$V = \frac{2.3mol \times 0.0821 \times 314K}{1.70atm}$$

3. Calculate

$$V = \frac{59.29}{1.70}$$

4. Answer

$$V = 34.88$$

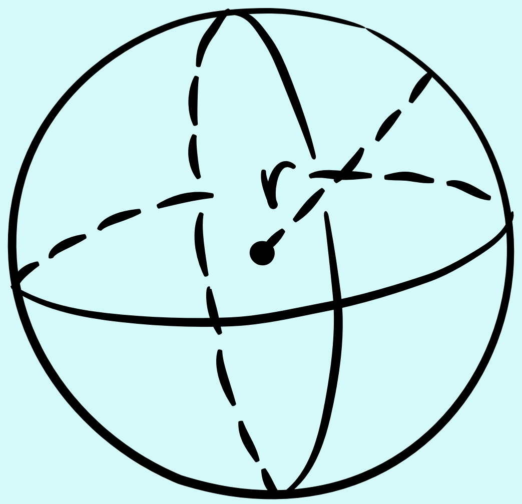
Therefore, there is 35L of gas

Question 2:

At a certain temperature, 3.24 moles of CO_2 gas at 2.15 atm take up a volume of 35.28L what is this temperature (in Celsius)?

$\frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

$y = mx + b$



$V = \frac{4}{3}\pi r^3$

ANSWER:

At a certain temperature, 3.24 moles of CO2 gas at 2.15 atm take up a volume of 35.28L what is this temperature (in Celsius)?

1. Ideal Gas Law

Equation: $PV = nRT$

- P = 2.15atm
- V = 35.28L
- n= 3.24mol
- R = Universal gas constant (0.0821 L·atm/(mol·K))
- T = ?

2. Rearrange equation $T = \frac{PV}{nP}$

$$T = \frac{2.15atm \times 35.28L}{3.24mol \times 0.0821}$$

3. Calculate

$$T = \frac{75.85}{0.266} = 285.1536K$$

4. Answer

$$T = 12^{\circ}C$$

Therefore, the temperature is 12°C

$\frac{\pm \sqrt{b^2 - 4a}}{2a}$

$= mx + b$

$V = \frac{4}{3}\pi r^3$

Why do we use Molar Mass

What is Molar Mass

- Molar mass is the mass of one mole of a substance (a mole is a large number, 6.022×10^{23}, of atoms or molecules). It is usually measured in grams per mole (g/mol).

Finding the Molar Mass of a Gas

- To find the molar mass of a gas, you need to know the mass of the gas in grams and how many moles of the gas you have. **This is provided in the first sheet of the CRFS exam.**

Converting Grams to Moles

- Grams to moles can be converted with this formula:
- Number of moles = Mass of the gas in grams / molar mass of the gas

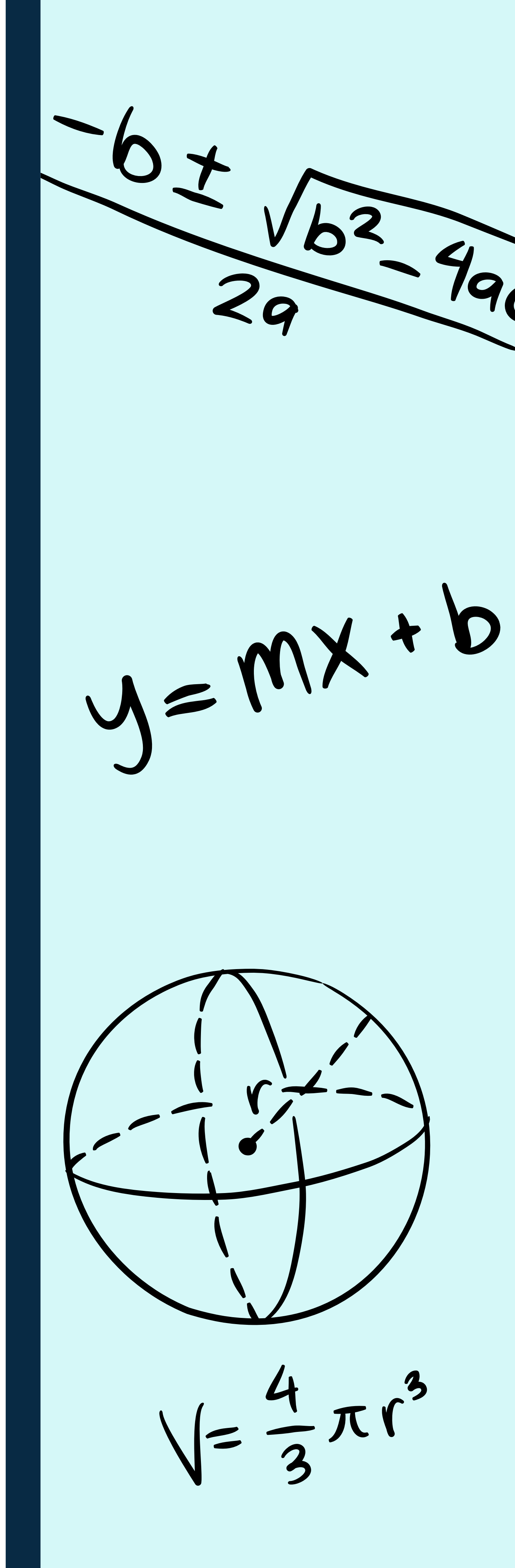
Example how many moles of O2 in 4g of O2?

$$\text{Moles of } O_2 = \frac{4 \text{ grams}}{32 \text{ g/mol}} = 0.125 \text{ moles}$$

- The molar mass of oxygen is about 32 g/mol.
- So, if you know the mass of a gas and its molar mass, you can easily find out how many moles of the gas you have!

QUESTION 3:

Calculate the volume that 12.5g of CO2 gas will occupy at a temperature of 40°C and a pressure of 1.20atm?



ANSWER:

Calculate the volume that 12.5g of CO2 gas will occupy at a temperature of 40°C and a pressure of 1.20atm?

1. Ideal Gas Law

Equation: $PV = nRT$

- P = 1.20atm
- V = ?
- n = ?
- R = Universal gas constant (0.0821 L·atm/(mol·K))
- T = 40°C

Gas Molar Masses

H_2	–	2.02	g mol^{-1}
He	–	4.00	g mol^{-1}
CH_4	–	16.04	g mol^{-1}
CO	–	28.01	g mol^{-1}
N_2	–	28.02	g mol^{-1}
O_2	–	32.00	g mol^{-1}
CO_2	–	44.01	g mol^{-1}
SF_6	–	146.06	g mol^{-1}

2. Convert everything

- 40°C = 313K
- CO2 = 12.0 + 2(16.0) ∴ CO2 = 44g/mol
- Number of moles = $\frac{12.5g}{44.0\frac{g}{mols}} = 0.284moles$

3. Rearrange equation

$$V = \frac{nRT}{P}$$

4. Calculate

$$V = \frac{0.284mol \times 0.0821 \times 313K}{1.20atm}$$

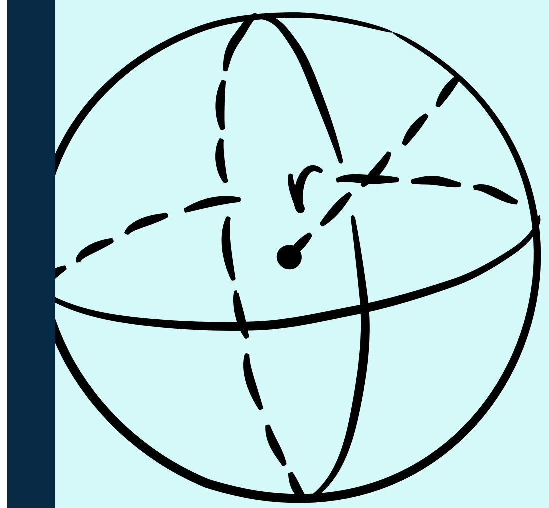
5. Answer

$$V = 6.08L$$

Therefore, the volume of gas occupied by 12.5g of CO2 at a temperature of 40°C and a pressure of 1.20atm is **6.08L**

$\frac{-b \pm \sqrt{b^2 - 4a}}{2a}$

$y = mx + b$



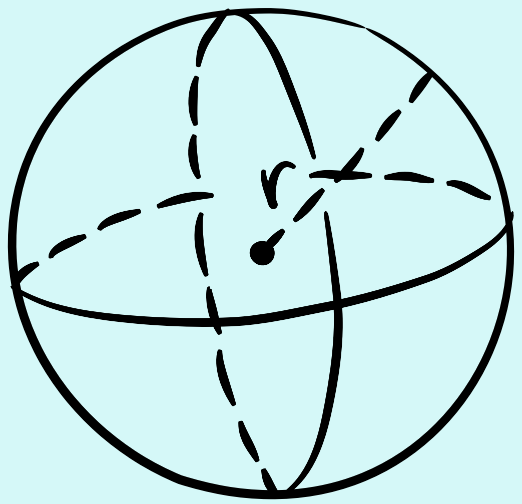
$V = \frac{4}{3}\pi r^3$

Question 4:

A certain gas has a molar mass of 28.0g/mol. How many grams of this gas would fit in a 3.00L container at 182kPa and 47.2°C?

$\frac{-b \pm \sqrt{b^2-4a}}{2a}$

$y = mx + b$



$V = \frac{4}{3}\pi r^3$

ANSWER:

A certain gas has a molar mass of 28.0g/mol. How many grams of this gas would fit in a 3.00L container at 182kPa and 47.2°C?

1. Ideal Gas Law

Equation: $PV = nRT$

- P = 182kPa
- V = 3.00L
- n = ?
- R = Universal gas constant (0.0821 L·atm/(mol·K))
- T = 47.2°C

2. Convert everything

- 182kPa = **1.80atm** (182/101.3) kPa to atm
- 47.2°C = **320.2K**

3. Rearrange equation

$$n = \frac{PV}{RT}$$

4. Calculate

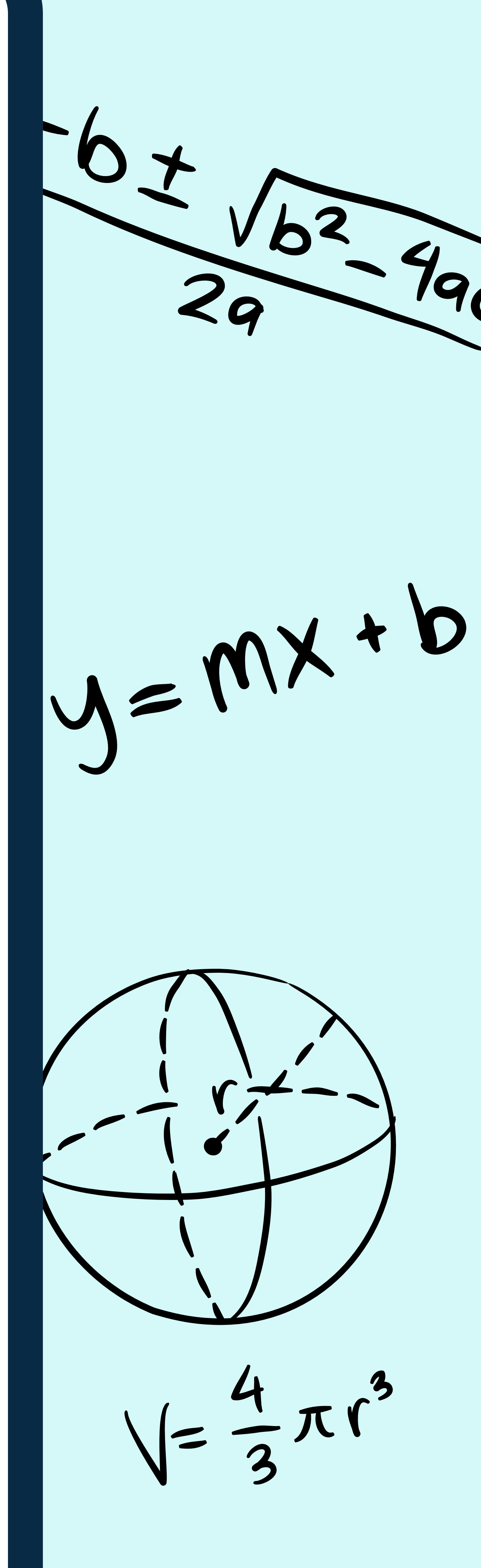
$$n = \frac{1.80atm \times 3.00L}{320.2K \times 0.0821}$$

5. Answer

$$n = 0.206moles$$

If the gas had 28.0g/mol and we have identified that the mass weighs 0.206 moles...

28.0g/moles * 0.206moles = 5.77grams



Why do we use density

What is density

- Density makes it easy to describe how much mass of a gas is present in a given volume

Simple calculations

- You can directly relate the physical properties of a gas (like its mass per volume) to the conditions it's under (pressure and temperature).

Real-world applications

- This is particularly useful in real-world applications where you want to know how heavy a certain volume of gas will be, like in weather forecasting or engineering.

If... $Density = \frac{mass\,(m)}{Volume\,(V)}$

and... $PV = nRT$

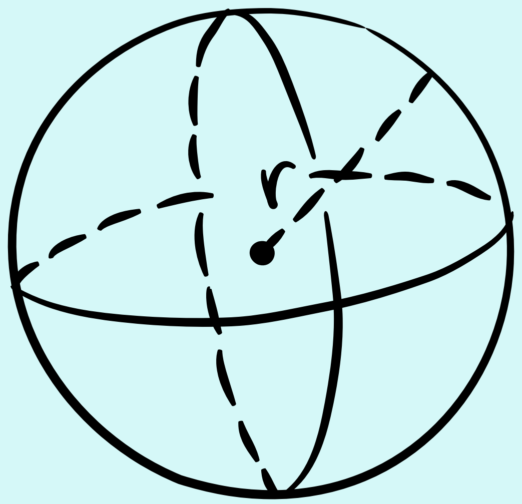
$$n = \frac{PV}{RT} \rightarrow \frac{m}{M} = \frac{PV}{RT} \rightarrow \frac{m}{V} = \frac{PM}{RT} \rightarrow D = \frac{PM}{RT}$$

Question 5:

What is the density in g/L of a sample of SO_2 at 932mmHg and 65°C?

$$\frac{-b \pm \sqrt{b^2-4ac}}{2a}$$

$$y = mx + b$$



$$V = \frac{4}{3}\pi r^3$$

Answer:

What is the density in g/L of a sample of SO2 at 932mmHg and 65°C?

1. Ideal Gas Law

Equation: $PV = nRT$

- P = 932mmHg
- M= Molar mass of SO2
- D= ?
- R= Universal gas constant (0.0821 L·atm/(mol·K))
- T = 65°C

2. Convert everything

- 65°C = 338K
- 932mmHg = 1.23mol
- Molar Mass of SO2 = 32.1+2(16.0) = 64.1g/mol

3. Rearrange equation

$$D = \frac{PM}{RT}$$

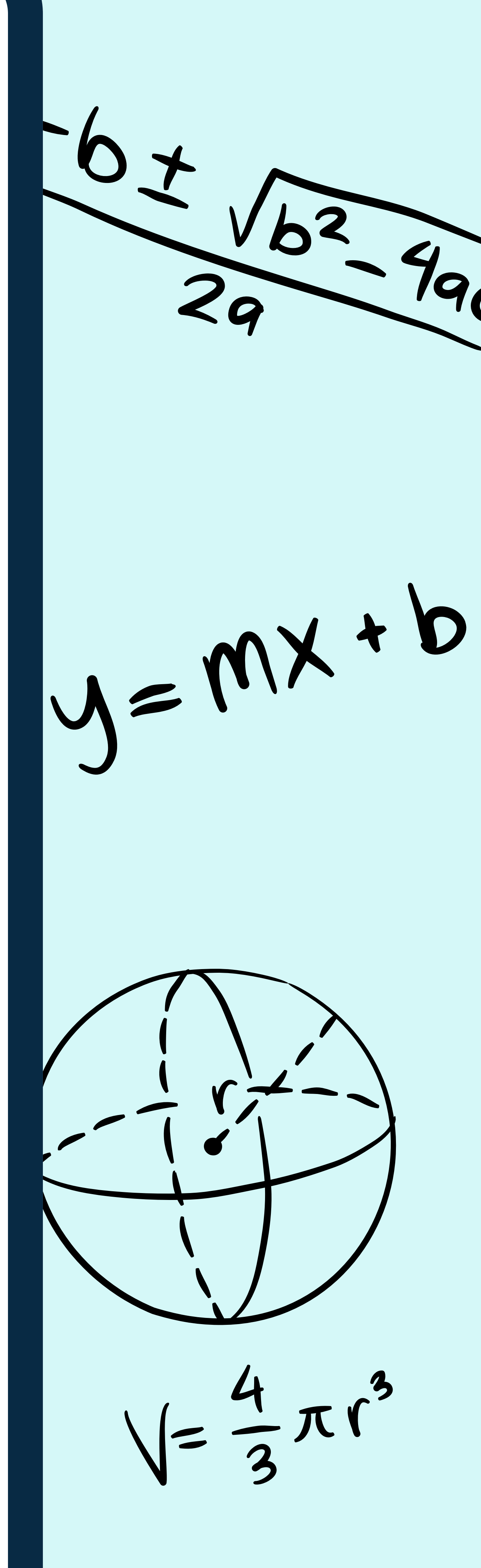
4. Calculate

$$D = \frac{1.23atm \times 64.1\frac{g}{mol}}{0.0821 \times 338}$$

5. Answer

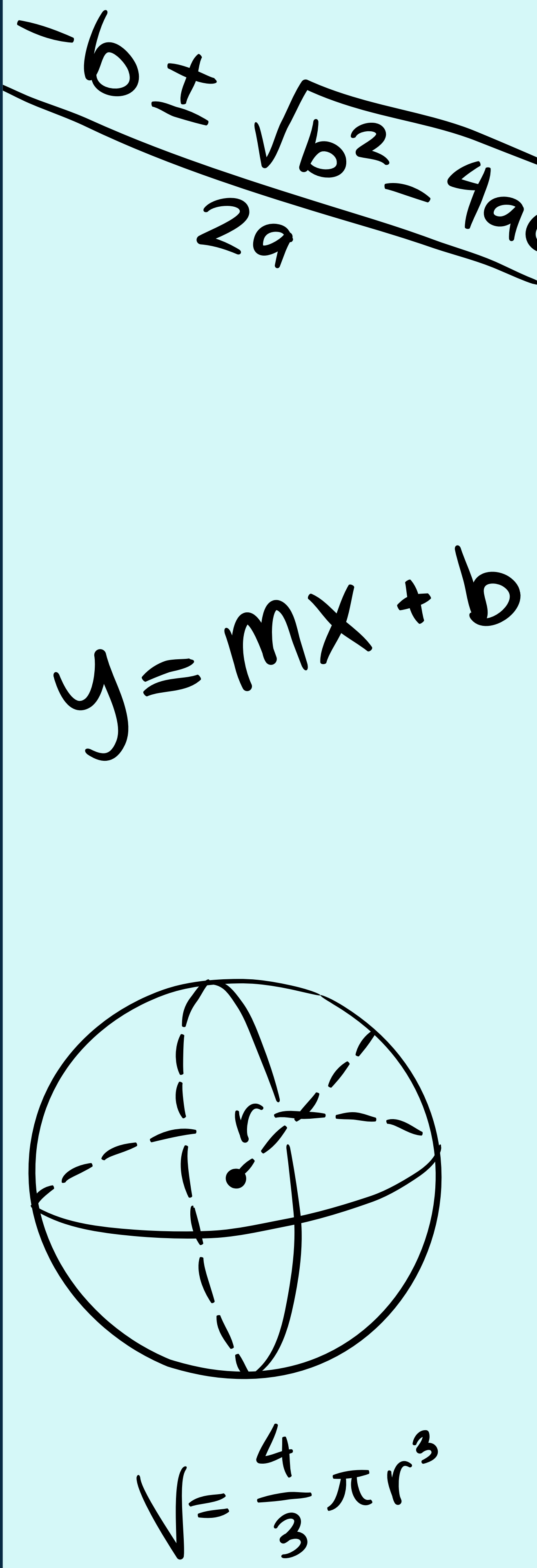
$$D = 2.84\frac{g}{L}$$

Therefore, the density of the gas of SO2 at 932mmHg and 65°C is **2.84g/L**



Question 6:

A gas has a density of 3.01g/L at STP what is its molar mass?



ANSWER:

A gas has a density of 3.01g/L what is its Molar mass

1. Ideal Gas Law

Equation: $PV = nRT$

- P = 1atm
- M= ?
- D= 3.01g/L
- R= Universal gas constant (0.0821 L·atm/(mol·K))
- T = 273K

2. Convert everything

- STP = 1atm & 0°C

3. Rearrange equation

$$D = \frac{PM}{RT}$$

$$M = \frac{RTD}{P}$$

4. Calculate

$$M = \frac{0.0821 \times 273K \times 3.01\frac{g}{mol}}{1}$$

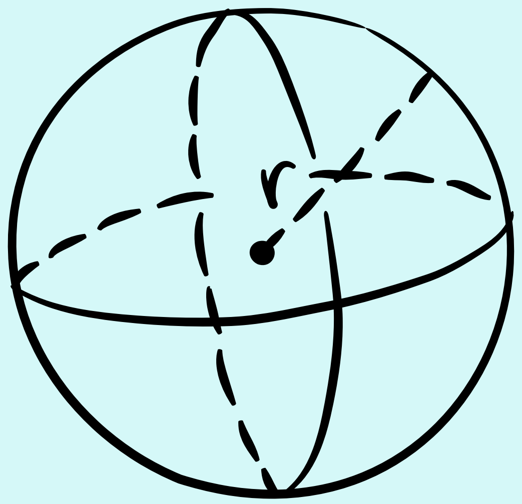
5. Answer

$$M = 67.5\frac{g}{mol}$$

Therefore, the Molar mass of a gas with a density of 3.01g/L is **67.5g/mol**

$$\frac{-b \pm \sqrt{b^2-4a}}{2a}$$

$$y = mx + b$$



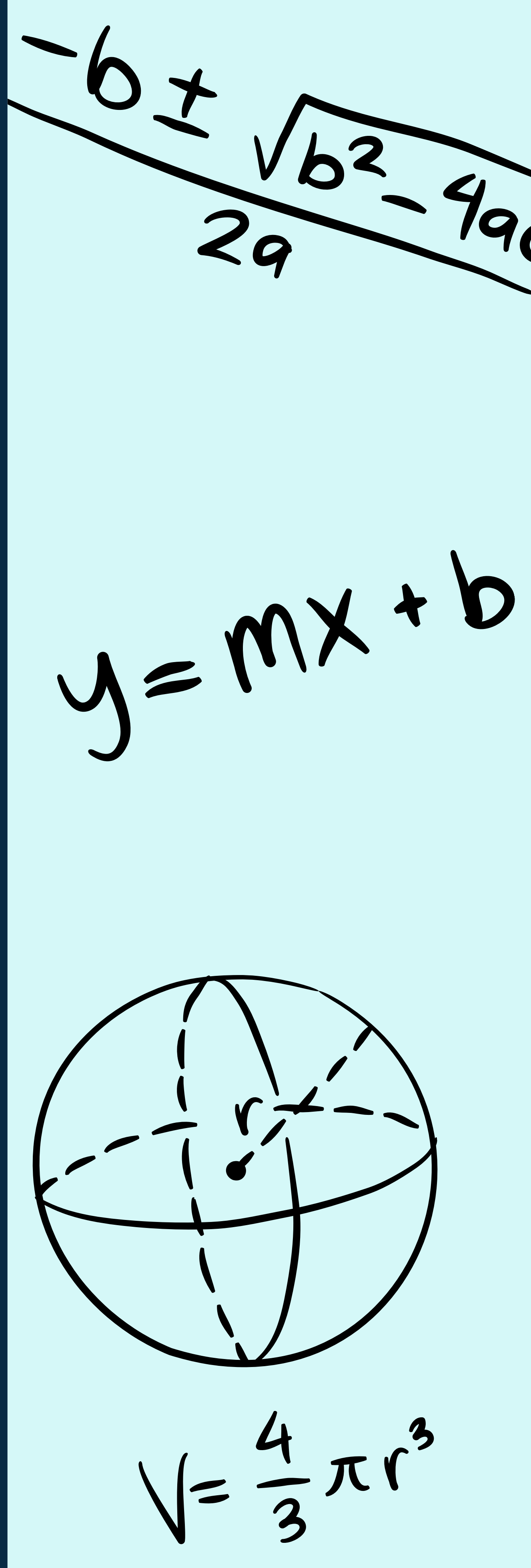
$$V = \frac{4}{3}\pi r^3$$

Question 7:

Subtopic(s) 3.07

CRFS question in the 2023 October exam

How much does 1L of air weigh?



Answer Part 1: How much does 1L of air weigh?

Gas Molar Masses

H_2	–	2.02	g mol^{-1}
He	–	4.00	g mol^{-1}
CH_4	–	16.04	g mol^{-1}
CO	–	28.01	g mol^{-1}
N_2	–	28.02	g mol^{-1}
O_2	–	32.00	g mol^{-1}
CO_2	–	44.01	g mol^{-1}
SF_6	–	146.06	g mol^{-1}

1. **Air composition and Molar Masses**

- Nitrogen (N_2): ~78.09% by volume
- Oxygen (O_2): ~20.95% by volume
- Argon (Ar): ~0.93% by volume
- Carbon Dioxide (CO_2): ~0.04% by volume

2. Calculate the Average Molar Mass

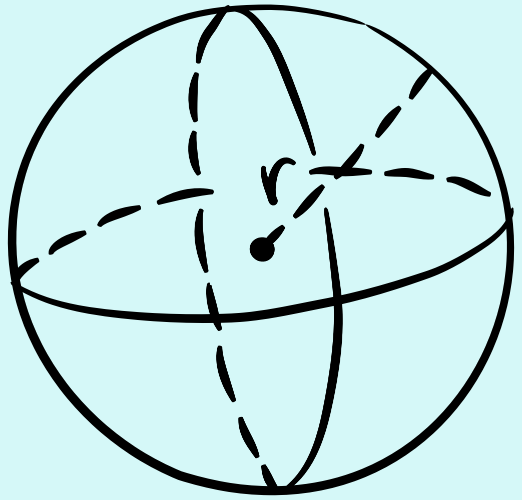
- (0.7809×28.97)+(0.2095×32.00)+(0.0093×39.95)+(0.0004×44.01)
- Average Molar Mass=(22.63)+(6.70)+(0.37)+(0.018)
- Average Molar Mass≈29.72 g/mol

3. Ideal gas law equation (PV=nRT)

- P = 1atm
- V = 1L
- n = ?
- R = Universal gas constant (0.0821 L·atm/(mol·K))
- T = 273K

$= \frac{-b \pm \sqrt{b^2 - 4a}}{2a}$

$y = mx + b$



$V = \frac{4}{3}\pi r^3$

Answer Part 2: How much does 1L of air weigh?

3.Ideal gas law equation (PV=nRT)

- P = 1atm
- V = 1L
- n = ?
- R = Universal gas constant (0.0821 L·atm/(mol·K))
- T = 273K

$$n = \frac{1atm \times 1L}{0.0821 \times 273.15K}$$

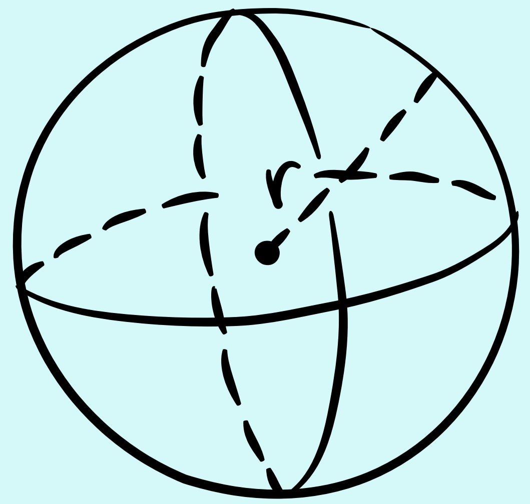
$$n = 0.04464$$

4. Calculate the Mass of 1 Liter of Air

- If Mass = n * Average molar mass
- Mass = 0.04464mol * 29.72g/mol
- Mass = 1.325grams

$$x = \frac{-b \pm \sqrt{b^2 - 4a}}{2a}$$

$$y = mx + b$$



$$V = \frac{4}{3}\pi r^3$$

Answer Part 3: How much does 1L of air weigh?

Alternatively we can calculate the weight of air with density

Step 1: Understand the Density of Air

- **Density of air at sea level and at 0°C (273.15 K):** The density of dry air is approximately **1.225 kg/m³**.

Step 2: Convert Volume from Liters to Cubic Meters

- 1 liter (L) is equal to 0.001 cubic meters (m³).

Step 3: Calculate the Mass of 1 Liter of Air

Using the formula:

$$\text{Mass} = \text{Density} \times \text{Volume}$$

Substitute the values:

$$\text{Mass} = 1.225\,\text{kg}/\text{m}^3 \times 0.001\,\text{m}^3$$

$$\text{Mass} = 0.001225\,\text{kg}$$

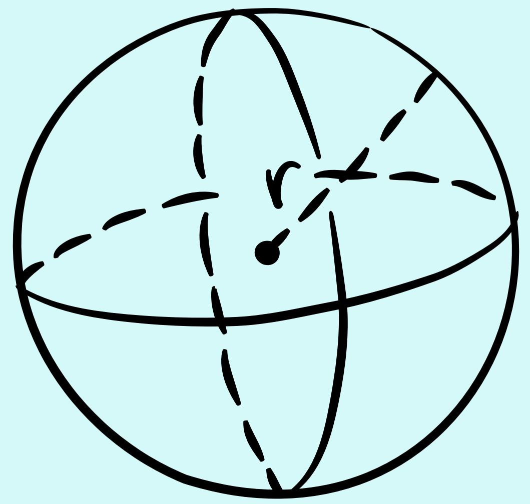
Step 4: Convert Mass to Grams

Since 1 kg = 1000 grams:

$$\text{Mass} = 0.001225\,\text{kg} \times 1000\,\text{g}/\text{kg} = 1.225\,\text{g}$$

$$x = \frac{-b \pm \sqrt{b^2 - 4a}}{2a}$$

$$y = mx + b$$



$$V = \frac{4}{3}\pi r^3$$