

PHYSICAL PHARMACY I

The Gaseous State The Gaseous State

BEHAVIOR OF GASES



By:
**Asst. Professor
Lena M. Thomas**



Contents

In this and the following lecture you will learn:

- ▶ **What is Gas and what are the properties of Gases ?**
- ▶ **Gas Laws and ideal gas equation**
 - ▶ How to calculate M. wt of gas using ideal gas equation?
- ▶ **Ideal and Real Gas** : what are the differences ?
- ▶ **Van der Waals modified gas equation** (for real gases)
- ▶ **Liquefaction of Gases:**
 - ▶ what does it mean and how to achieve it ?
 - ▶ Aerosols as an example for gas liquefaction

What is Gas ?

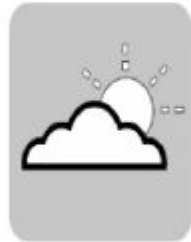
Gas is one of the fundamental states of matter.

A pure gas may be made up of

- **Individual atoms** (inert gases, e.g He gas)
- **Elemental molecules** (made from one type of atom, e.g., N₂ gas)
- **Compound molecules** (made from a variety of atoms, e.g., CO₂ gas)

The Atmosphere is all around us

→ an "ocean" of gases mixed together



Composition

nitrogen (N₂)..... ~78%

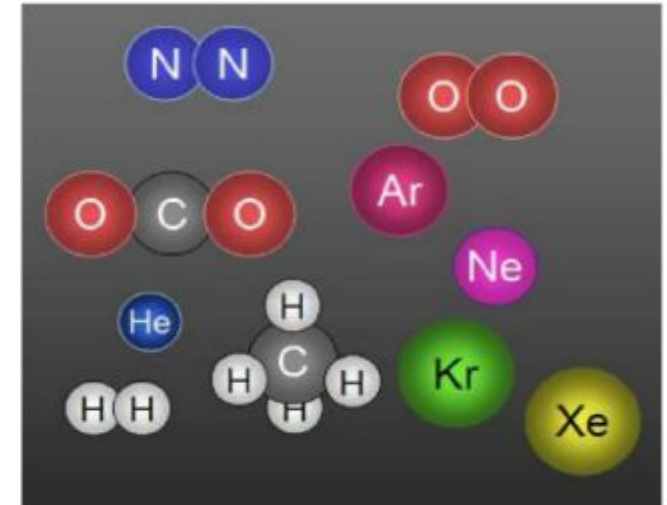
oxygen (O_2).....~21%

argon (Ar).....~1%

carbon dioxide (CO₂)... ~0.04%

water vapor (H_2O)..... ~0.1%

Trace amounts of:
He, Ne, Rn, SO₂,
CH₄, N_xO_x, etc.



Elements that exist as gases at 25°C and 1 atmosphere

[illegible]

Properties of Gases

Compared to other states of matter:

Gases have **negligible intermolecular forces between its particles**, are **separated by large distances** and gas particles **move at high velocity**.

Gases show the **most uniform behavior** irrespective of the nature of the gas and **have the following properties:**

- (i) They expand indefinitely and fill up the whole vessel in which they are placed.
- (ii) They can be compressed by the application of pressure.
- (iii) They can be liquefied by cooling and applying pressure.
- (iv) They possess low densities under ordinary conditions of temperature and pressure.
- (v) They intermix spontaneously, i.e., they show the property of *diffusion*.
- (vi) They exert pressure on the walls of the vessel in which they are contained.

Measurable Properties of Gases

Variables describing a gas

P = Pressure of gas in atmospheres (atm)

V = Volume of gas in liters (L)

n = Number of moles of gas in moles (mol)

T = Temperature of gas in Kelvin (K)

R = Ideal Gas Constant $0.082057 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$

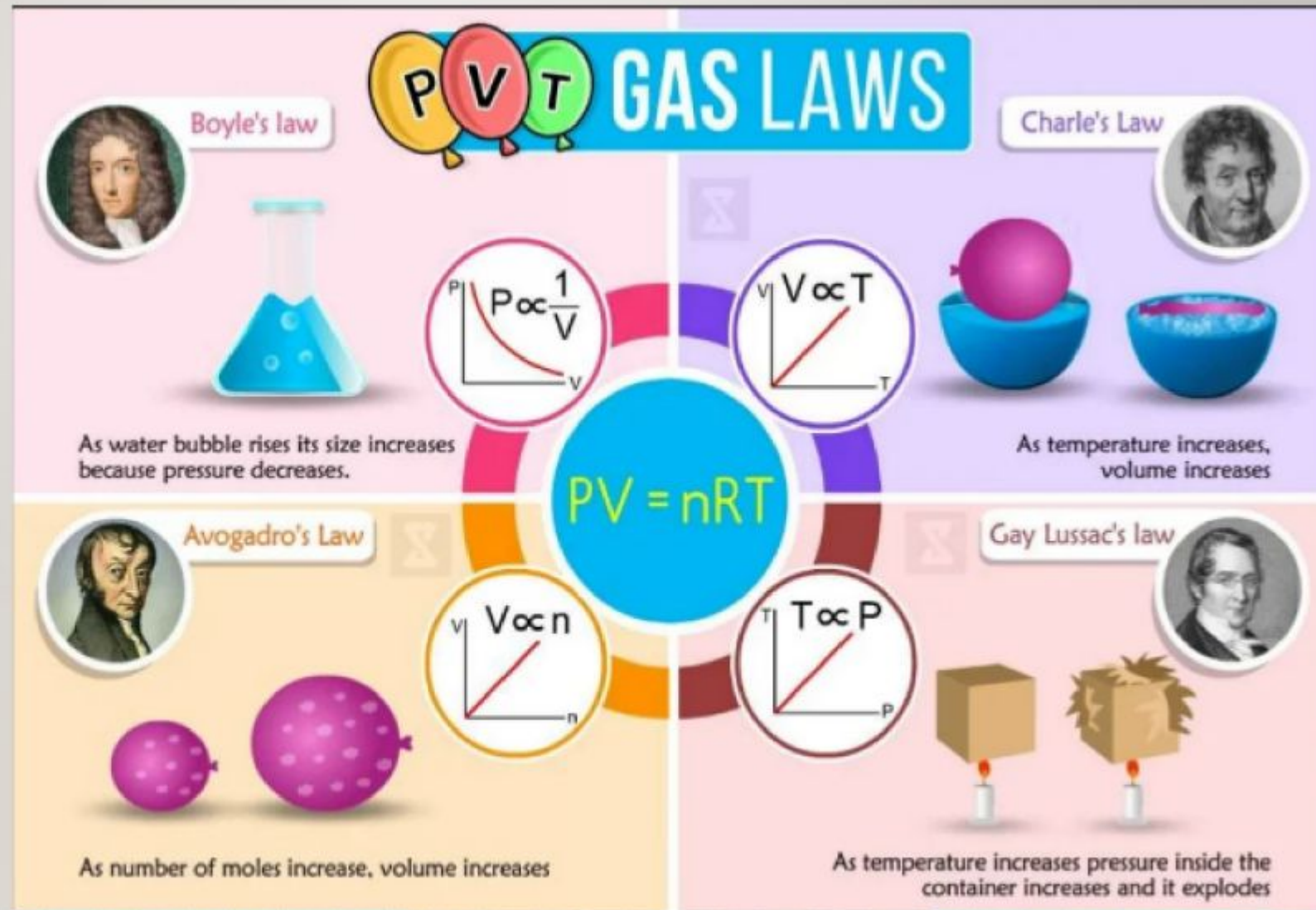
Units of pressure: since pressure is force /area
so: dyne/cm^2 ,

Or: $1 \text{ atm} = 760 \text{ torr} = 760 \text{ mmHg} = 1 \times 10^5 \text{ Pa}$
 $= 14.69 \text{ psi} = 101.325 \text{ kPa}$

$1 \text{ L} = 1000 \text{ mL} = 1000 \text{ cm}^3$

zero $^{\circ}\text{C} = 273 \text{ Kelvin}$,
so: degree in K = $^{\circ}\text{C} + 273$

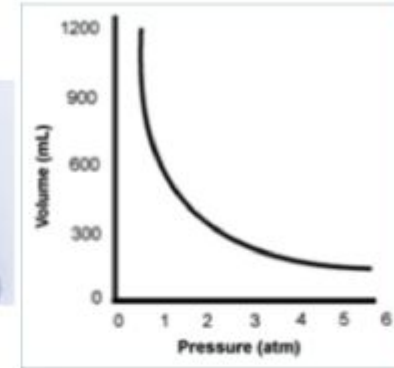
An overview on Gas Laws



1.

Boyle's Law

The **volume** of a given mass of a gas is **inversely proportional** to its **pressure** with **temperature** remaining constant.



Mathematically, if V is the volume of a gas at pressure P , and if temperature is kept constant, then:

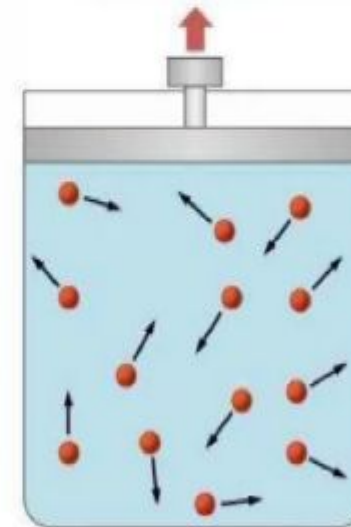
$$P \propto \frac{1}{V}$$

or

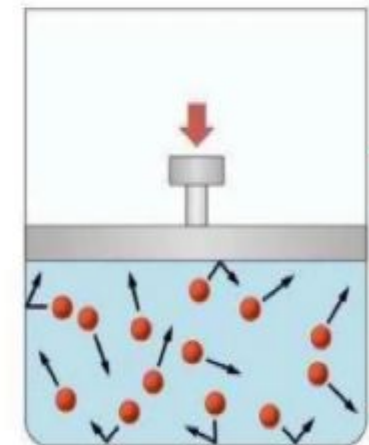
$$PV = k$$

In other words, if at constant temperature, V_1 is the volume of a gas at pressure P_1 and on changing the pressure to P_2 , the volume changes to V_2 , then

$$P_1 \times V_1 = P_2 \times V_2$$



As volume increases,
pressure decreases



As volume decreases,
pressure increases



Examples

A sample of gas in a cylinder has a volume of **720 ml** under pressure of **0.375 atm**.

What will be the volume occupied by gas under **1 atm** when the temperature remains constant ?

$$P_1 \times V_1 = P_2 \times V_2$$

$$0.375 \times 720 = 1 \times V_2$$

$$V_2 = 270 \text{ ml}$$

A gas occupies 100. mL at 150. kPa and 25°C. Find its volume at 200. kPa and 25°C.

$$P_1 \times V_1 = P_2 \times V_2$$

$$150 \times 100 = 200 \times V_2$$

$$V_2 = 75 \text{ ml}$$

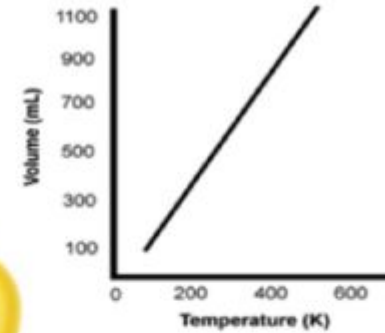
Exercise:

- A gas has a volume of 500 milliliters when a pressure equivalent to 760 millimeters of mercury is exerted upon it. Calculate the volume if the pressure is reduced to 730 millimeters.

2.

Charles Law

The **volume** of a given mass of a gas is **directly proportional** with **temperature** at constant **pressure**



Mathematically, if V is the volume of a gas and T is the temperature, and pressure is kept constant, then:

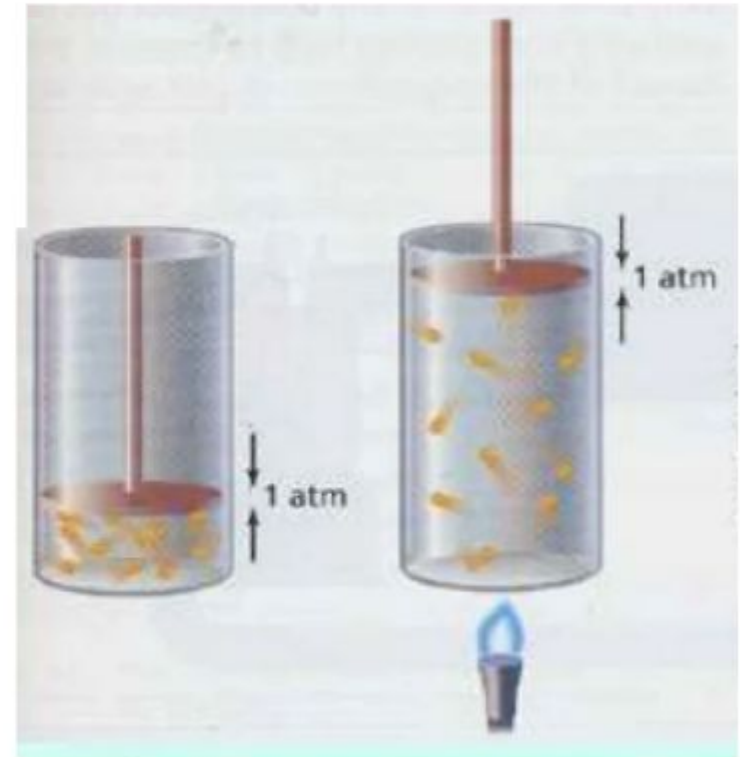
$$V \propto T$$

$$V = kT$$

Or:

$$\frac{V}{T} = k \text{ or } \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Temperature must be in Kelvin units



Example

A mass of gas has a volume of 150 ml at temperature of 25 ° C. What will be the volume of gas at 45 ° C if pressure was held constant ?

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

$$V_1 = 150 \text{ ml}, \quad T_1 = 25 \text{ C} = 25 + 273 = 298 \text{ Kelvin}$$

$$V_2 = ? \quad T_2 = 45 \text{ C} = 45 + 273 = 318 \text{ Kelvin}$$

$$\frac{150}{298} = \frac{V_2}{318}$$

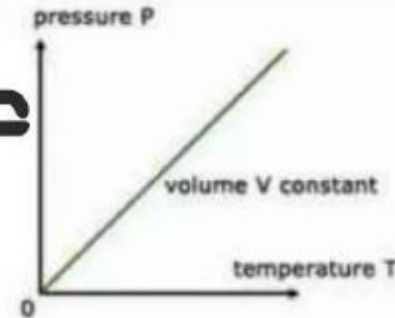
$$V_2 = 160 \text{ mL}$$

Exercise

A gas occupies 473 cm³ at 36°C and 1 atm. Find its volume at 94°C and 1 atm.

3. Gay-Lussac's Law

The **pressure** of a given mass of a gas is **directly proportional** with **temperature** at constant **volume**

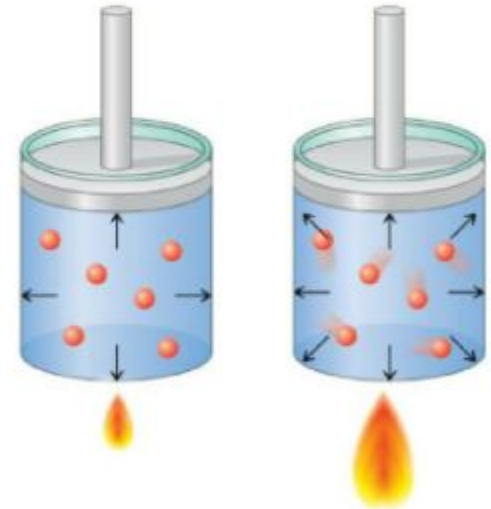


Mathematically, if P is the pressure of gas and T is the temperature, and volume is kept constant, then:

$$P \propto T$$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Temperature
must be in
Kelvin units



Remember: increasing the temperature gets the particles moving faster causing an increase in the number and energy involved in collisions; thus an increase in pressure

■ Should you throw an aerosol can into a fire? What could happen?

4.

Avogadro's Law

Avogadro's number (6.023×10^{23}) : Is the number of molecules (regardless their identity whether is it a gas, liquid, or solid) present in 1 mole (n) of substance.



Under similar conditions of temperature and pressure,
equal volumes of all gases contain the same number of molecules.

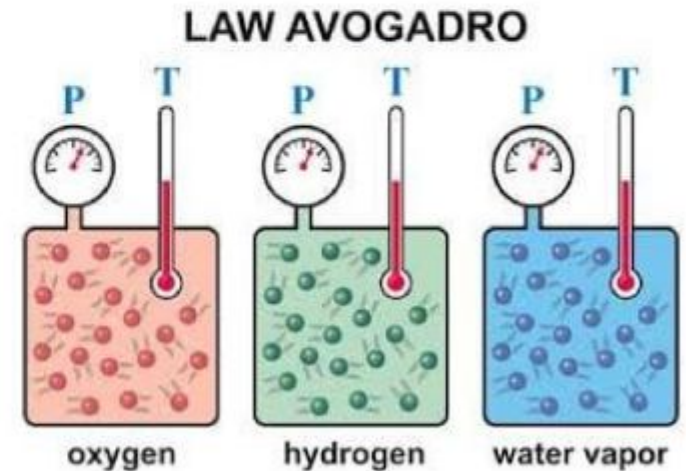
The **volume (V)** of a gas is directly proportional to the **number of moles (n)** present in the gas if **temperature (T)** and **pressure (P)** are kept constant.



$$V \propto n$$

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

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$



Example

Consider 2 samples of N_2 . Sample 1 contains 1.5 mol N_2 with a volume of 36.7 L at 25°C and 1 atm. Sample 2 has a volume of 16.5 L at 25°C and 1 atm. How many moles of N_2 are in Sample 2?

$$V_1 = 36.7 \text{ L}$$

$$n_1 = 1.5 \text{ mol}$$

$$V_2 = 16.5 \text{ L}$$

$$n_2 = ?$$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

$$\frac{36.7 \text{ L}}{1.5 \text{ mol}} = \frac{16.5 \text{ L}}{n_2}$$
$$n_2 = 0.67 \text{ mol}$$

A sample with 0.15 moles of gas has a volume of 2.5 L. What is the volume if the sample is increased to 0.55 moles?

$$n_1 = 0.15 \text{ mol}$$

$$V_1 = 2.5 \text{ L}$$

$$n_2 = 0.55 \text{ mol}$$

$$V_2 = ?$$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

$$\frac{2.5 \text{ L}}{0.15 \text{ mol}} = \frac{V_2}{0.55 \text{ mol}}$$
$$V_2 = \frac{(0.55 \text{ mol})(2.5 \text{ L})}{(0.15 \text{ mol})}$$
$$V_2 = 9.2 \text{ L}$$

If 0.75 mole of helium gas occupies a volume of 1.5 L, what volume (L) will 1.2 moles of helium occupy at the same temperature and pressure?

Data:

$$n_1 = 0.75 \text{ mol}$$

$$V_1 = 1.5 \text{ L}$$

$$n_2 = 1.2 \text{ mol}$$

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

$$\frac{n_1}{V_1} = \frac{n_2}{V_2}$$

$$V_2 = \frac{V_1 \times n_2}{n_1}$$

$$V_2 = \frac{1.5 \text{ L} \times 1.2 \cancel{\text{mol}}}{0.75 \cancel{\text{mol}}}$$

$$V_2 = 2.4 \text{ L}$$

The Combined Gas Law Equation (the Ideal Gas Equation)

The previously mentioned laws and equations (summarized in the box to the right) may be combined to give a **general equation**, called the **gas equation** (or more correctly, **ideal gas equation** since real gases do not follow this equation)

$$\begin{aligned} V &\propto 1/P \text{ (Boyle's law)} \\ V &\propto T \text{ (Charles's law)} \\ V &\propto n \text{ (Avogadro's law)} \end{aligned}$$

Mathematically, ideal gas equation is given by:

$$\frac{PV}{T} = R$$

Under one set of conditions, although P, V and T change, the ratio PV/T is constant and can be represented as:

$$PV = RT \quad R \text{ is a constant value ; the gas constant}$$

This equation is correct for 1 mole (i.e., g molecular weight) of gas: for n moles it becomes:

$$PV = nRT$$

P = pressure

V = volume

n = amount of substance

R = ideal gas constant

T = temperature

$$1 \text{ atm} = 760 \text{ mmHg}$$

$$n = \text{wt (gm)}/\text{M.wt (molar)}$$

$$= 0.082 \text{ L. atm/mole. K}$$

Ideal gas law for 2 set of condition is:

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

Note that in this form the gas constant R does not appear (omitted from both sides since it's the same)

Example

What is the volume of one mole of an ideal gas at 25° C and 780 mmHg ?

$$PV = nRT$$

$$P = 1 \text{ atm} = 760 \text{ mmHg} : \text{so} : 780 \text{ mmHg} = 1.026 \text{ atm}$$

$$n = 1 \text{ mole}$$

$$V = ?$$

$$R = 0.082 \text{ L.atm/mole.K}$$

$$T = 25 + 273 = 298 \text{ Kelvin}$$

$$1.026 \times V = 1 \times 0.082 \times 298$$

$$V = 74.6 \text{ L}$$

Exercises

Calculate the volume in liters of 1 mole of nitrogen gas at 400 atm and 0°C, using the ideal gas equation. Give the answer in milliliters of N₂.

Example

What is the volume of a weather balloon rising 2 miles into the upper atmosphere at atmospheric pressure of $8.77 \times 10^{-3} \text{ atm}$, and temperature of -44.7°C , if you know that it has initial volume of 2.5 L at ground level at 1 atm pressure and 24°C .

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$P_1 = 1 \text{ atm},$$

$$P_2 = 8.77 \times 10^{-3} \text{ atm}$$

$$T_1 = 24 + 273 = 297, \quad T_2 = -44.7 + 273 = 228.3$$

$$V_1 = 2.5 \text{ L}$$

$$V_2 = ??$$

$$\frac{1 \times 2.5}{297} = \frac{8.77 \times 10^{-3} \cdot V_2}{228.3}$$

$$V_2 = 219.6 \text{ L}$$

How do you calculate molecular weight of a gas?

To find the molecular mass of a gas, we can assume it behaves ideally and use the ideal gas law

$$PV = nRT$$

We can modify this law in by replacing the number of moles of gas (n) by its equivalent g/M, in which g is the mass (m) and M is molecular weight of the gas:

$$PV = \frac{g}{M}RT$$

Thus, to find the molecular weight of a gas, we can measure its volume, mass, pressure and temperature.

Example:

If 0.5 g of drug in the vapor state occupies 100 ml at 120 °C and 1 atm pressure. What is approximate M.wt (assuming that the gas behaves ideally)

Answer:

$$P = 1 \text{ atm} = 760 \text{ mmHg}$$

$$Wt = 0.5 \text{ gm}$$

$$V = 100 \text{ mL} = 0.1 \text{ L}$$

$$R = 0.082 \text{ L.atm/mole.K}$$

$$T = 120 + 273 = 393 \text{ Kelvin}$$

$$PV = nRT$$

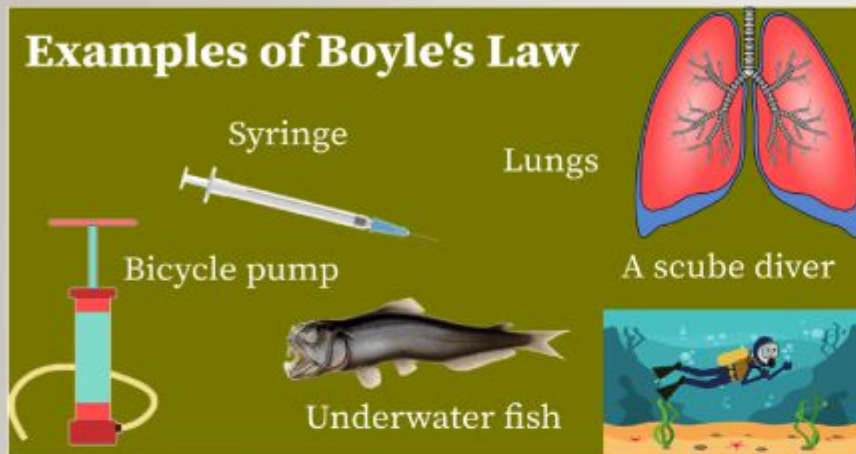
$$PV = \frac{wt}{M.wt} RT$$

$$M.wt = \frac{wt \times R \times T}{P.V}$$

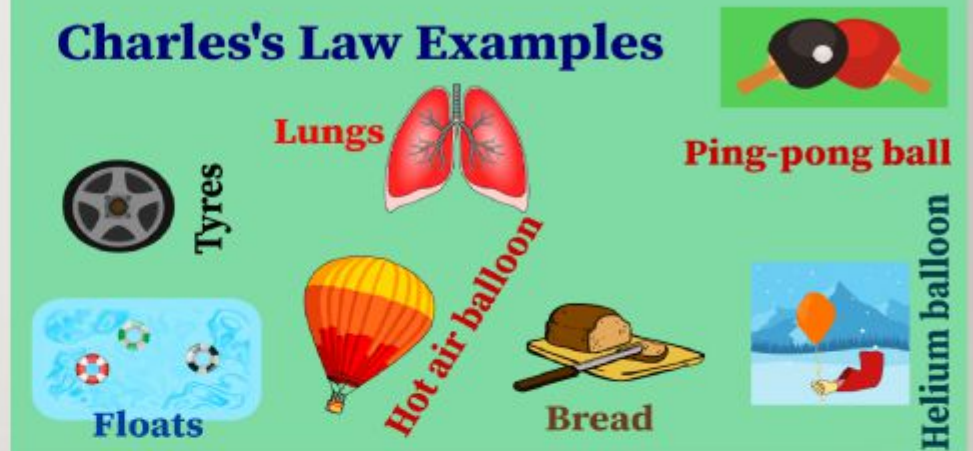
$$= \frac{0.5 \text{ g} \times 0.082 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 393 \text{ K}}{1 \text{ atm} \times 0.1 \text{ L}}$$

$$= 161.13 \text{ g/mole}$$

Examples of Boyle's Law



Charles's Law Examples



Real-life Examples of Gay-Lussac's Law

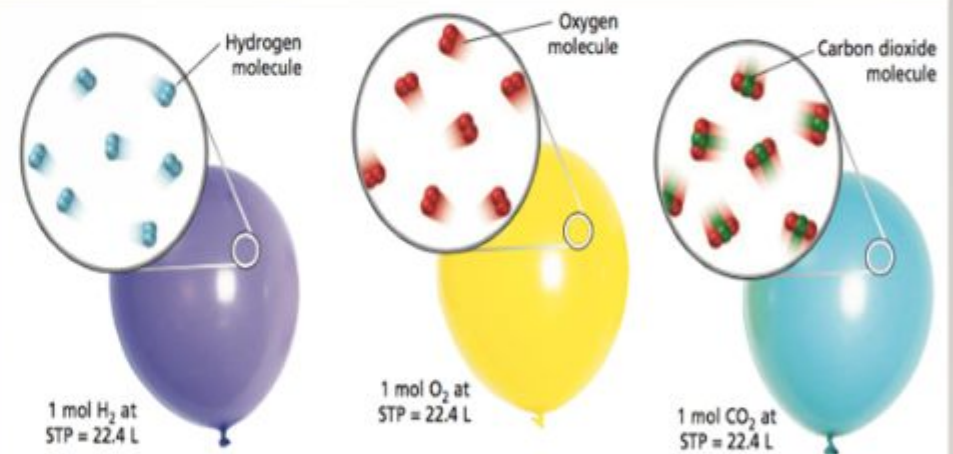


Fig. 11 from Modern Chemistry, Raymon Davis et al, HRW

At the same temperature and pressure, balloons of equal volume have equal numbers of molecules, regardless of which gas they contain.

Standard temperature and pressure (STP)

- As the volume of a given mass of gas differs with temperature and pressure, the STP conditions are used as a reference point to compare gases.
- STP refers to the nominal conditions of pressure and temperature which are the at sea level atmosphere.



Standard Temperature & Pressure

0°C 273 K
- OR -
1 atm 101.325 kPa
 760 mm Hg

Standard molar volume (also called, molar volume) is the volume of 1 mole of gas at STP.

For ideal gas :

$$PV = nRT$$

$$V = nRT/P$$

$$V = (1)(0.082)(273)/(1)$$

$$V = 22.40 \text{ L}$$



Diagram showing three balloons of different colors (yellow for Ne, blue for O₂, and purple for CO₂) tied together, illustrating that 1 mole of any gas occupies the same volume at STP.

Volume:	22.4 L	22.4 L	22.4 L
Pressure:	1 atm	1 atm	1 atm
Temperature:	273 K	273 K	273 K
Quantity:	1 mole	1 mole	1 mole
Mass:	40.0 g	32.0 g	28.0 g