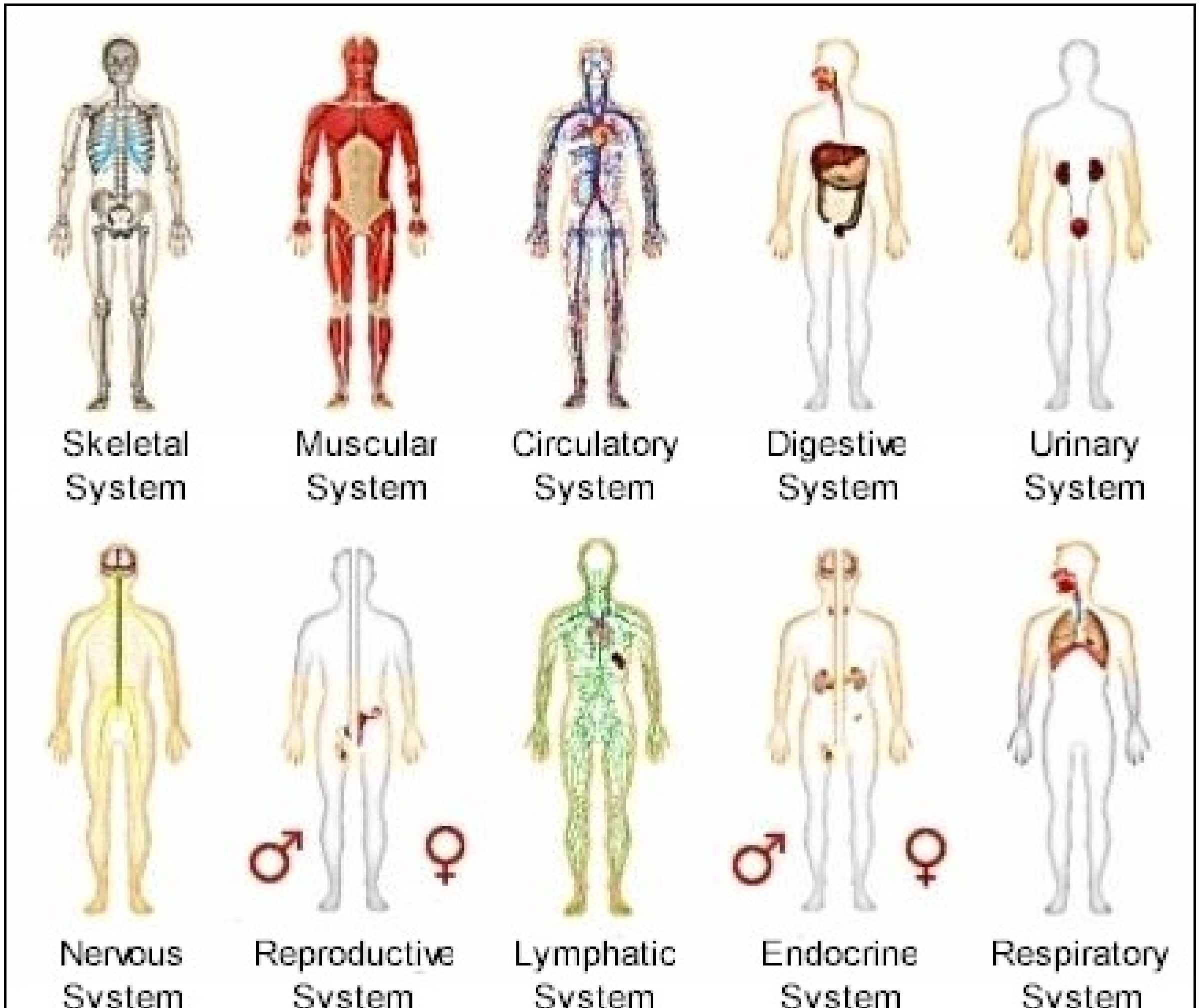
HUMAN PHYSIOLOGY (인체 생리학)



Human Physiology

physio (nature)   + logia (science)

자연의 성질이나 그 기능을 연구하는 학문

- 인체를 구성하는 세포와 조직이 가지고 있는 기능을 연구하여 생체에서 일어나는 여러 process, activity 및 function을 통한 인체의 부분 또는 전체의 기능을 나타내게 하는 기전 (mechanism)에 대해 연구하는 학문

생리학(Physiology)이란?

- 생명체 기능들에 관해 연구하는 학문
- 생명체 기능들을 설명하기 위한 2가지 접근법
  - 생명체 공정의 **목표**에 대해 강조(**왜?**)
    - 생체가 필요로 하는 것을 충족하기 위해 설명
  - 기능의 **기작(mechanism)**에 대해 강조(**어떻게?**)
    - 원인과 효과 과정들로 설명
      - 생명체를 하나의 기계로 봄

- 생리학은 해부학과 밀접한 관련
- **해부학(anatomy)**: 생명체 구조에 관한 학문
- 생리학적 기작은 구조 체계를 통해 가능
- 생명체 구조-기능의 연관(예, 심장은 혈액을 받아들이고 펌핑)
- 기능들을 이해하기 위해 해부학적 지식 필수

분류

- Non living body
- Living body : biological science
  - Morphology
  - Function

Must Have Characteristics:

- Able to move
- Able to reproduced
- Able to grow

Might Have Characteristics:

- Be used as food by other organism
- Have leaves, roots, and stems
- Fly, Walk
- Go to school

Word/Topic/Concept

Living Things

Examples

- Insects
- Plants
- Birds
- Worms
- Snails
- Weeds

Nonexamples

- Rocks
- Steel
- Air
- Water
- Sound
- Fire

생명체의 구성 수준

생명체 구성의 계층적 구성 수준.

Basic Physical Chemistry

Gas Laws: Overview

- The three fundamental gas laws discover the relationship of pressure, temperature, volume and amount of gas.

- Boyle's Law tells us that the *volume* of gas increases as the *pressure* decreases.

- Charles Law tells us that the *volume* of gas increases as the *temperature* increases.

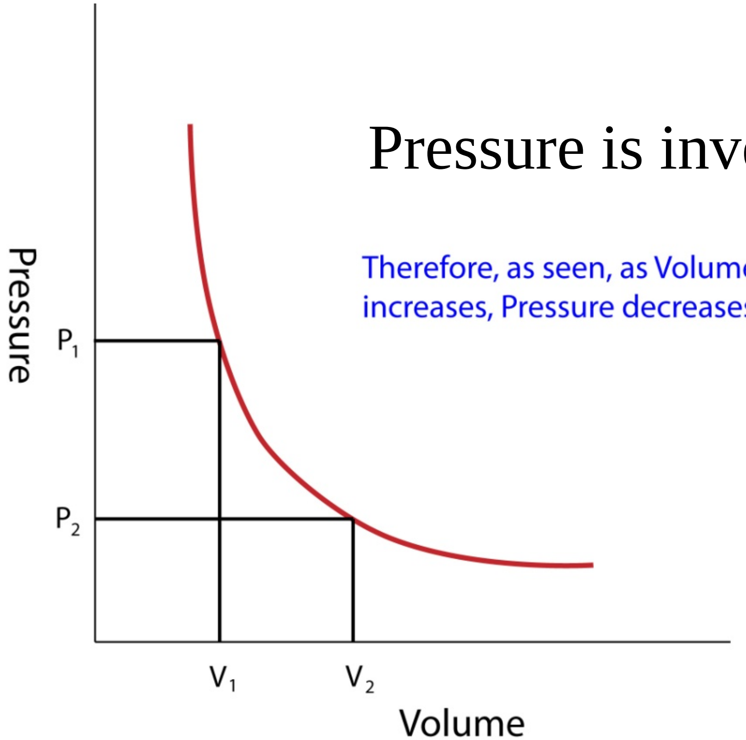
- Avogadro's Law tell us that the *volume* of gas increases as the *amount of gas* increases.

- The ideal gas law is the combination of the three simple gas laws.

Boyle's Law

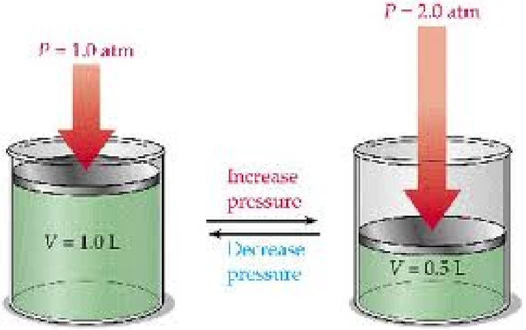
In 1662, Robert Boyle discovered the correlation between *Pressure (P)* and *Volume (V)* (assuming *Temperature(T)* and *Amount of Gas(n)* remain constant):

$$P \propto 1/V \qquad PV = constant.$$



Pressure is inversely proportional to Volume

Therefore, as seen, as Volume increases, Pressure decreases



Example: 1.1

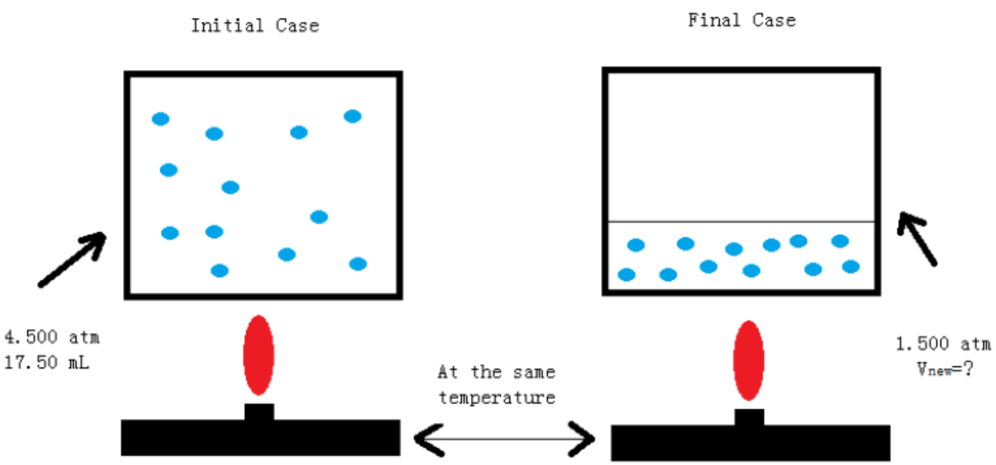
A 17.50mL sample of gas is at 4.500 atm. What will be the volume if the pressure becomes 1.500 atm, with a fixed amount of gas and temperature?

Solution:

$$V_2 = \frac{P_1 . V_1}{P_2}$$

$$= \frac{4.500atm.17.50mL}{1.500atm}$$

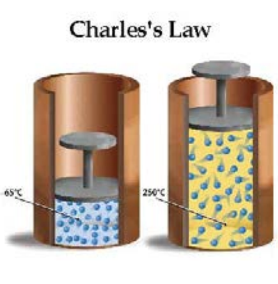
$$= 52.50mL$$



Charles' Law

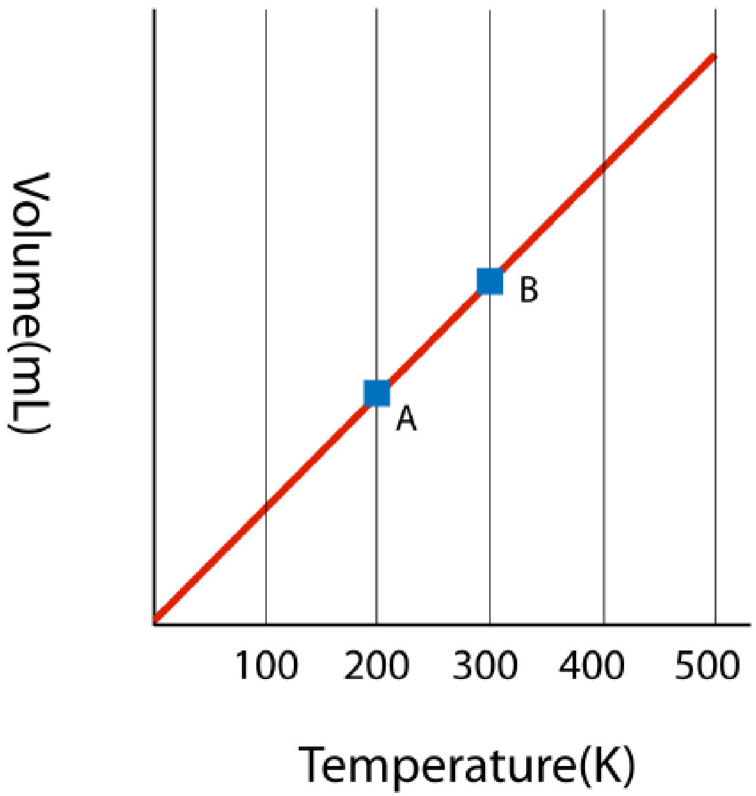
Charles's Law

THE RELATIONSHIP BETWEEN TEMPERATURE AND VOLUME

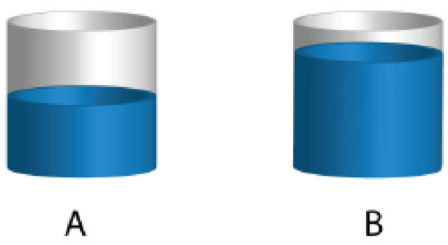


In 1787, French physicists Jacques Charles, discovered the correlation between *Temperature(T)* and *Volume(V)* (assuming *Pressure(P)* and *Amount of Gas(n)* remain constant):

$$V \propto T \qquad V_1T_2 = V_2T_1$$



As Temperature increases, Volume increases(linear graph)



Example: 1.2 A sample of Carbon dioxide in a pump has volume of 20.5mL and it is at 40.0 ^oC. When the amount of gas and pressure remain constant, find the new volume of Carbon dioxide in the pump if temperature is increased to 65.0 ^oC.

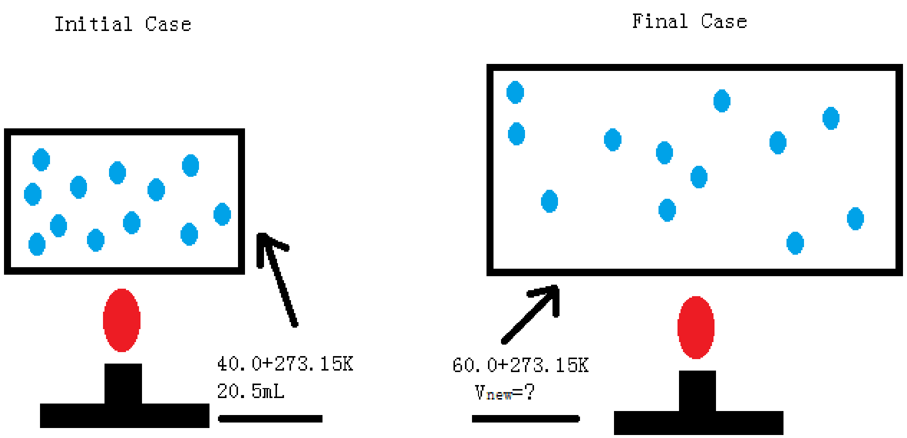


Figure 1.2

Solution:

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

$$= \frac{20.5mL.(60+273.15K)}{40+273.15K}$$

$$= 22.1mL$$

Avogadro's Law

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

-A principle stated in 1811 by the Italian chemist Amadeo Avogadro (1776 -1856) that equal volumes of gases at the same temperature and pressure contain the same number of molecules regardless of their chemical nature and physical properties.

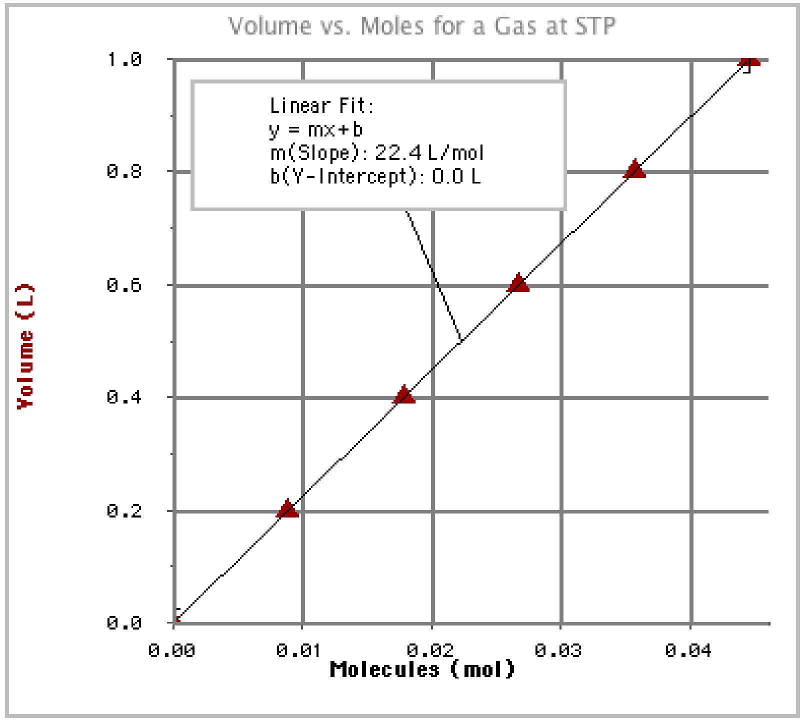
Avogadro's Law

In 1811, Amedeo Avogadro fixed Gay-Lussac's issue in finding the correlation between the *Amount of gas(n) and Volume(V) (assuming Temperature(T) and Pressure(P) remain constant):*

$$V \propto n \qquad V = zn \qquad (P,\ T = \text{constant})$$

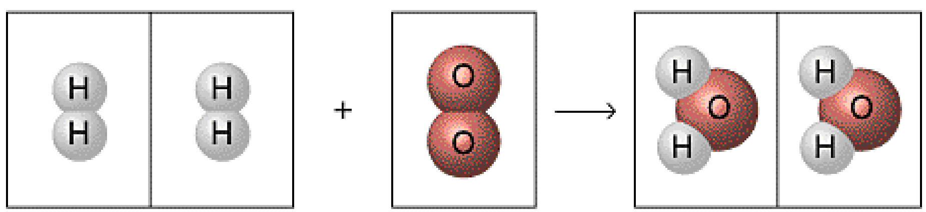
(where z is a constant depending on Pressure and Temperature)

$$\frac{P_1}{n_1} = z = \frac{P_2}{n_2}$$



Observation:	Two volumes hydrogen	+	One volume oxygen	$\longrightarrow$	Two volumes water vapor

Explanation:



Equation:	$2H_2(g)$	+	$O_2(g)$	$\longrightarrow$	$2H_2O(g)$

Example: A 3.80 g of oxygen gas in a pump has volume of 150mL. constant temperature and pressure. If 1.20g of oxygen gas is added into the pump. What will be the new volume of oxygen gas in the pump if temperature and pressure held constant?

Solution:

V_1=150mL

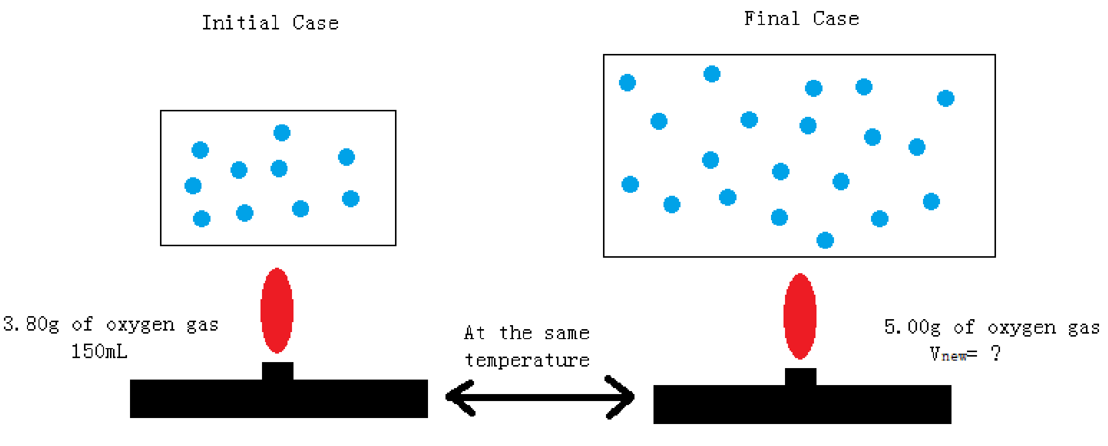
$n_1 = \frac{m_1}{M_{oxygen gas}}$

$N_2 = \frac{m_2}{M_{oxygen gas}}$

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

$$= \frac{150mL . \frac{5.00g}{32.0g.mol^{-1}}}{\frac{3.80g}{32.0g.mol^{-1}}}$$

$$= 197ml$$



Ideal Gas Law

Boyle's Law

$PV = k$

Charles's Law

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

P and *V* change
n, *R*, *T* are constant

Ideal Gas Law

$PV = nRT$

T and *V* change
P, *n*, *R* are constant

P, *V*, and *T* change
n and *R* are constant

Combined Gas Law

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

R = gas constant = 8.3145 Joules · mol^{-1} · K^{-1} (SI Unit)
= 0.082057 L · atm·K^{-1} · mol^{-1}

Example

At 655mm Hg and 25.0 ^{o}C, a sample of Chlorine gas has volume of 750mL. How many moles of Chlorine gas at this condition?

Solution:

P=655mm Hg

T=25+273.15K

V=750mL=0.75L

n=?

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

$$= \frac{655mmHg.\frac{1atm}{760mmHg}.0.75L}{0.082057L.atm.mol^{-}1.K^{-}1.(25+273.15K)}$$

$$= 0.026mol$$

Evaluation of the Gas Constant, R

You can get the numerical value of gas constant, R, from the ideal gas equation, PV=nRT.

At standard temperature and pressure, where temperature is 0°C, or 273.15K, pressure is at 1 atm, and with a volume of 22.4140L,

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

$$= \frac{1atm.22.4140L}{1mol.273.15K}$$

=0.082057 L atm mol^{-1} K^{-1}

Practice Problems

1. If 4L of H2 gas at 1.43 atm is at standard temperature, and the pressure were to increase by a factor of 2/3, what is the final volume of the H2 gas? (Hint: Boyle's Law)
2. If 1.25L of gas exists at 35oC with a constant pressure of .70 atm in a cylindrical block and the volume were to be multiplied by a factor of 3/5, what is the new temperature of the gas? (Hint: Charles's Law)
3. A ballon with 4.00g of Helium gas has a volume of 500mL. When the temperature and pressure remain constant.

What will be the new volume of Helium in the ballon if another 4.00g of Helium is added into the ballon? (Hint: Avogadro's Law)

Solution

3. 1000mL or 1L

Using Avogadro's Law to solve this problem, you can switch the equation into $V_2 = \frac{n_1 . V_2}{n_2}$

However, you need to convert grams of Helium gas into moles.

$n_1 = \frac{4.00g}{4.00g/mol} = 1$ mol   Similarily, n_2=2 mol

$$V_2 = \frac{n_2 . V_2}{n_1} = \frac{2mol . 500mL}{1mol}$$

Solution

1. 2.40L

In order to solve this question you need to use Boyle's Law:
Keeping the key variables in mind, temperature and the amount of gas is constant and therefore can be put aside,
the only ones necessary are:

1. Initial Pressure: 1.43 atm
2. Initial Volume: 4 L
3. Final Pressure: 1.43x1.67 = 2.39
4. Final Volume(unknown): V2

Plugging these values into the equation you get:
V2=(1.43atm x 4 L)/(2.39atm) = 2.38 L

2. 184.89 K

In order to solve this question you need to use Charles's Law:

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

Once again keep the key variables in mind. The pressure remained constant and since the amount of gas is not
mentioned, we assume it remains constant. Otherwise the key variables are:

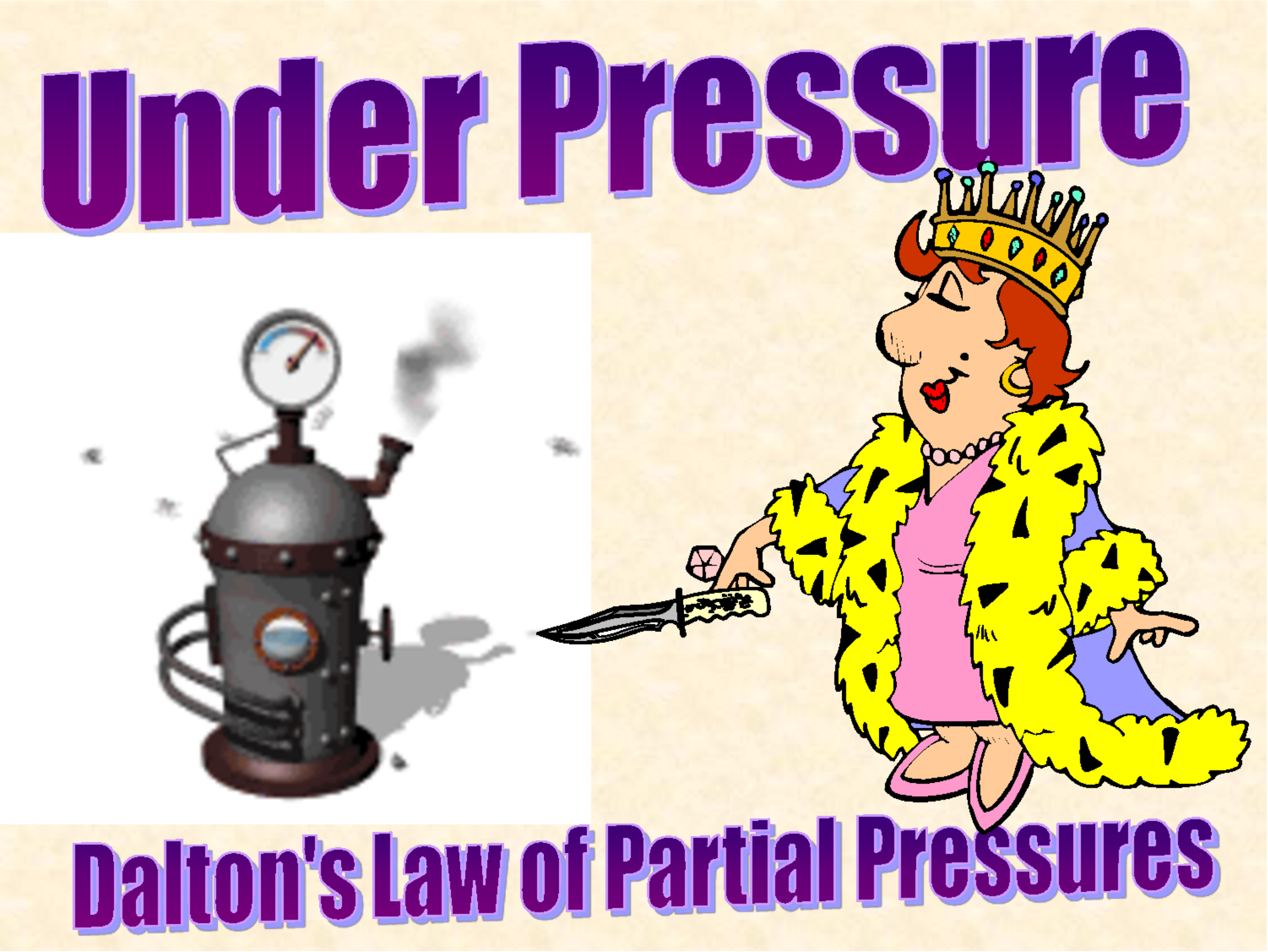
1. Initial Volume: 1.25 L
2. Initial Temperature: 35oC + 273.15 = 308.15K
3. Final Volume: 1.25L*3/5 = .75 L
4. Final Temperature: T2

Since we need to solve for the final temperature you can rearrange Charles's:

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

Once you plug in the numbers, you get: T2=(308.15 K x .75 L)/(1.25 L) = 184.89 K

Dalton's Law of Partial Pressures

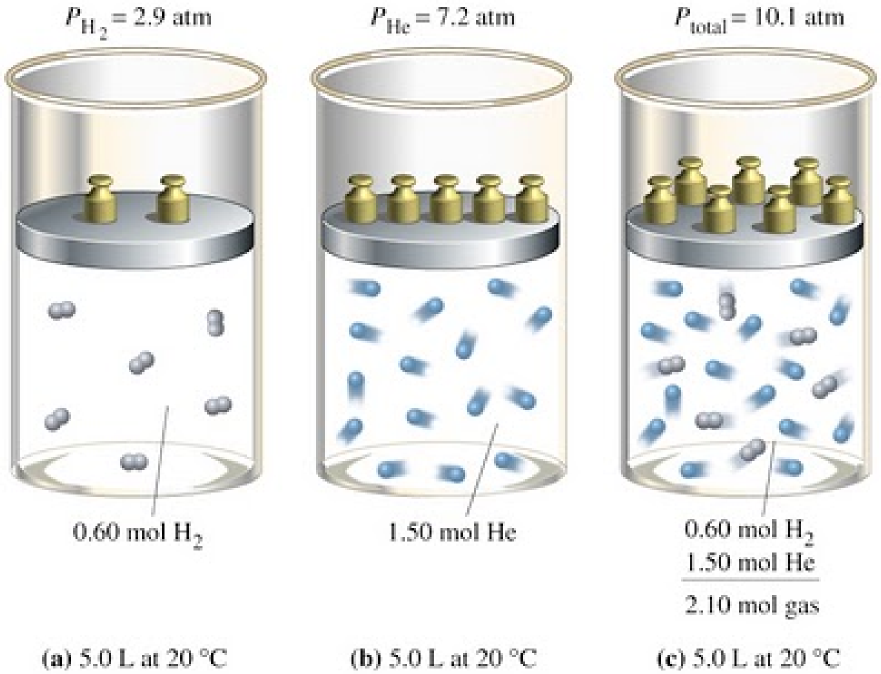


- Many gases are, like air, homogeneous mixtures of several gases.
- Because the gas particles do not interact in any way other than bouncing off each other, each gas particle has exactly the same influence on the gas law properties (P, V, T, and n) of the gas.
- One mole of nitrogen will exert the same pressure as a mixture of a tenth of a mole each of ten different gases under the same conditions of temperature and volume, assuming, of course, that the gases all behave like ideal gases.

In a mixture of gases, the total pressure is the sum of the partial pressures of all the gases present.

Dalton's Law of Partial Pressure:

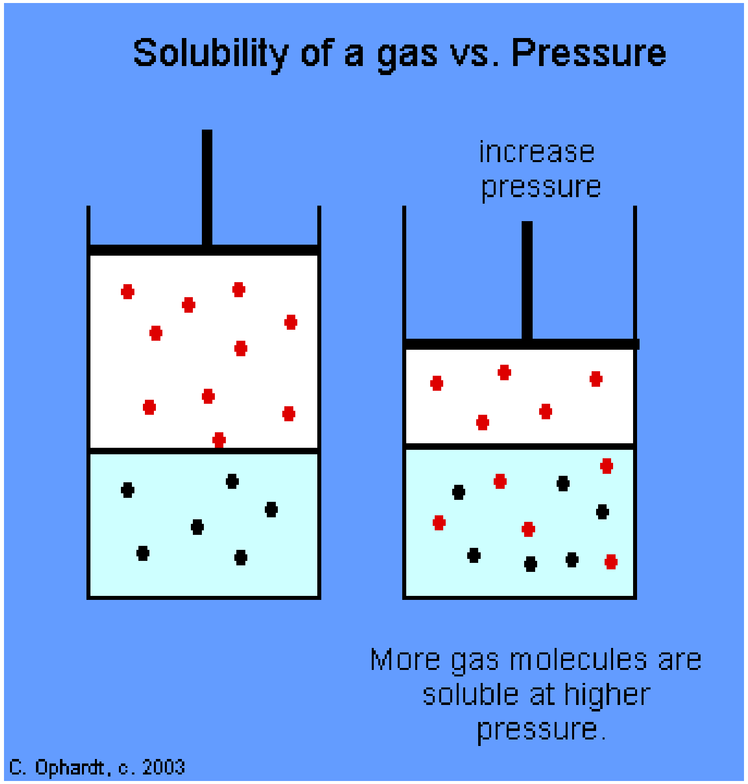
$$\text{Pressure}_{\text{Total}} = \text{Pressure}_1 + \text{Pressure}_2 \ldots \text{Pressure}_n$$



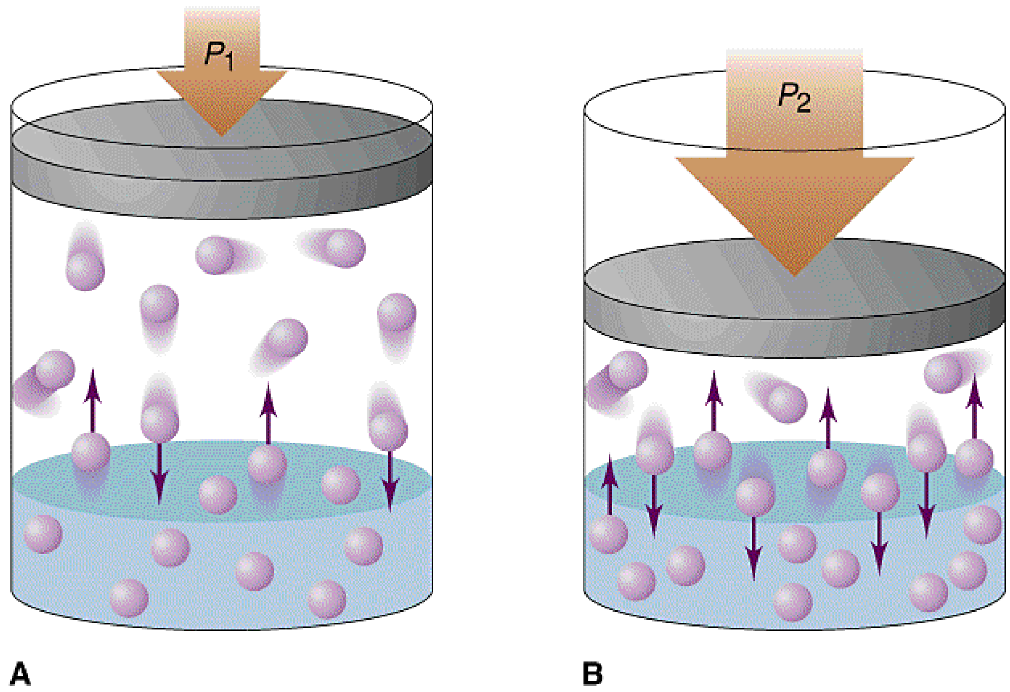
Henry's Law:

P=KC

Where P=pressure, C is the concentration of the gas and K is Henry's law constant.



This law tells the diver that at higher pressures our bodies will absorb more gases. The deeper the dive, the greater the amount of nitrogen absorbed into the body. Therefore the greater the depth the greater the risk of decompression illness.



$$\frac{C_1}{P_1} = \frac{C_2}{P_2}$$

The initial condition has concentration C_1 and gas partial pressure P_1. The second condition has concentration C_2 and gas partial pressure P_2.

Example:

What is the predicted concentration of dissolved oxygen, if the partial pressure for oxygen is 56 mm Hg? The concentration of dissolved oxygen is 0.44 g / 100 ml solution. The partial pressure of oxygen is 150 mm Hg.

Solution:

$$\frac{C_1}{P_1} = \frac{C_2}{P_2}$$

$P_1 = 150$ mm Hg $C_1 = 0.44$ g O_2 /100 ml solution

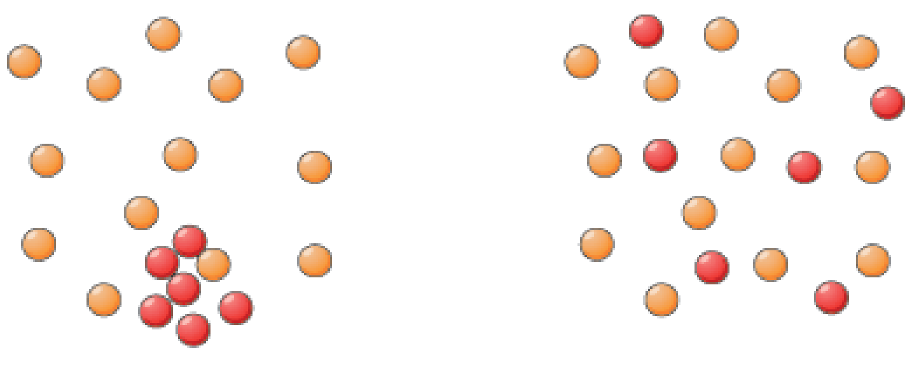
$P_2 = 56$ mm Hg $C_2 = ?$

$C_2 = 0.15$ g O_2 / 100 ml solution

Diffusion (확산)

Diffusion in *gases*

- When chemicals, like the smell of perfume or burning toast, are let loose in a room, the particles mix with the air particles.
- The particles of smelly gas are free to move quickly in all directions. They eventually spread through the whole room. This is called **diffusion**.
- Diffusion in gases is quick because the particles in a gas move quickly. It happens even faster in hot gases



Before diffusion After diffusion

Diffusion in *liquids*

- Diffusion can also happen in liquids. This is because the particles in liquids can move around each other, which means that eventually they are evenly mixed.
- For example if you drop a little bit of paint into a jar of water the colour will spread slowly through the water. This is by diffusion.
- Diffusion in liquids is slower than diffusion in gases because the particles in a liquid move more slowly.

DIFFUSION: THE PHENOMENON

Fick's Law of Diffusion

$$\text{Rate of Diffusion} = \frac{\text{Conc. Grad.} * \text{SA} * \text{Diffusion Coef.}}{\text{Membrane Thickness}}$$

SA = Surface Area

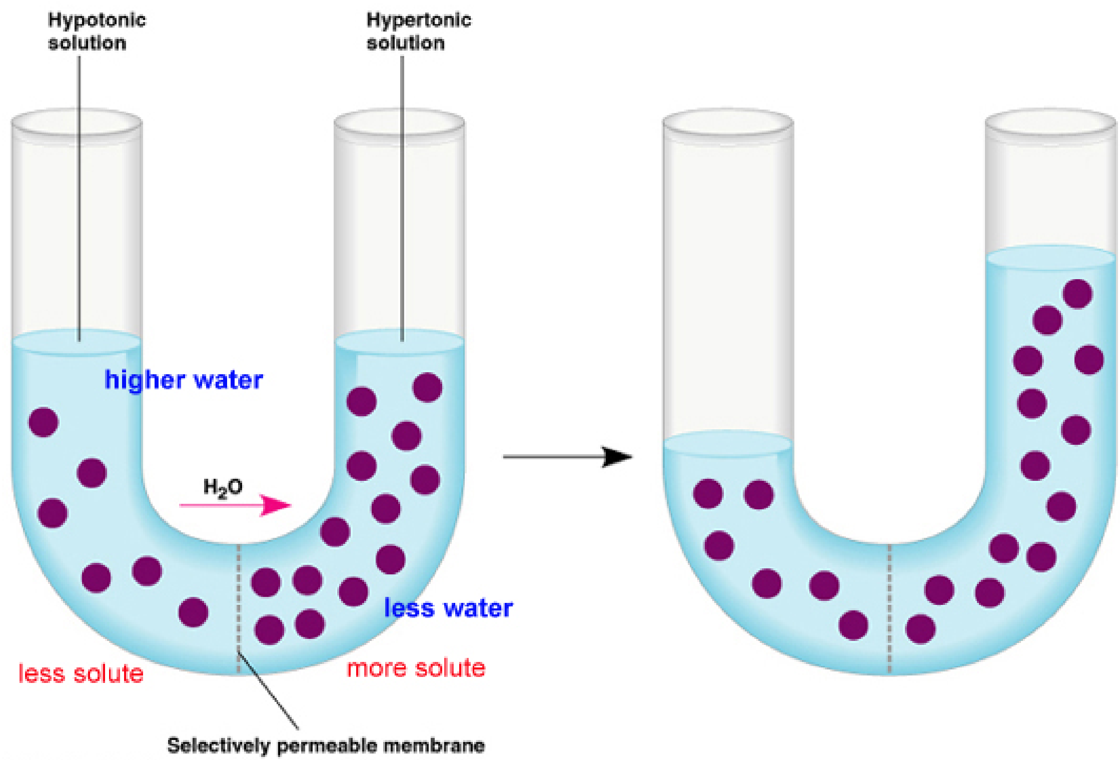
$$\text{Diffusion Coef.} = \frac{\text{Permeability}}{\text{SQRT (MW)}}$$

MW = Molecular Weight

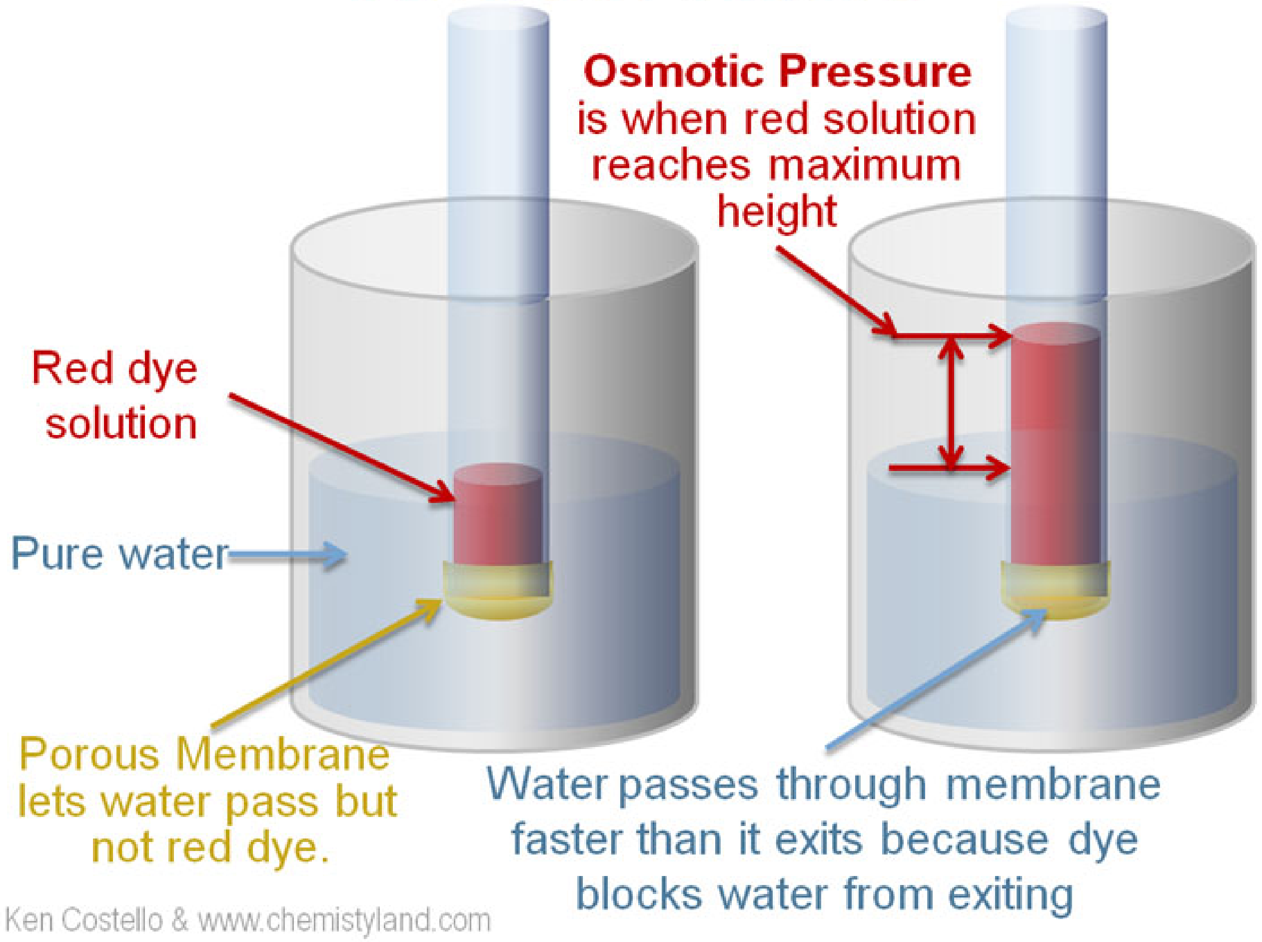
Example: K^+ has a lower MW and is 30X more permeable than Na^+.

Osmosis and Osmotic pressure

- Osmosis is the process by which water moves through a selectively permeable membrane.
- It is a special case of diffusion: it involves the diffusion of a solvent (like water), rather than the diffusion of substances dissolved in the solvent.



Osmotic Pressure



Osmotic pressure = pressure that has to be exerted on a solution to prevent it from gaining water when separated from pure water by an ideal selectively permeable membrane (it is positive).

$$\text{Osmotic pressure} = n \times \left(\frac{C}{M}\right) \times R \times T$$

n= number of particles into which the substance dissociates

C = concentration in g/L

M= molecular weight of the molecules

R= universal gas constant, which is 0.082

T= absolute temperature