

Topic:	GASES
Objective:	FK_06_01
<p>Given a mass of a gas and specific conditions the student must be capable of doing the following:</p> <ul style="list-style-type: none"> • calculate the amount of moles • calculate the value of pressure, volume or temperature • calculate the volumen when the conditions change (keeping the mass constant) 	

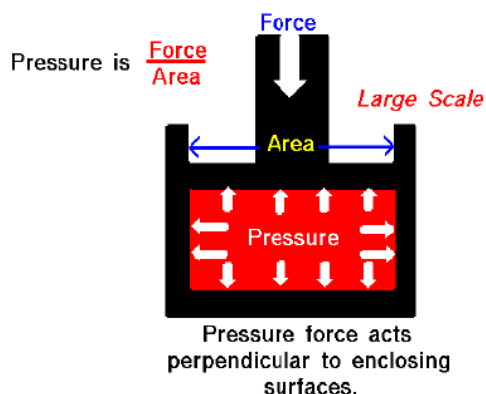
Pressure

- Pressure (P) is the force (F) per unit of area that acts on a given area (A):

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

- The SI unit for force is the Newton (kg.m/s²) and for pressure is the Pascal (Pa):

$$1 \text{ Pascal} = \frac{1 \text{ Newton}}{\text{m}^2}$$

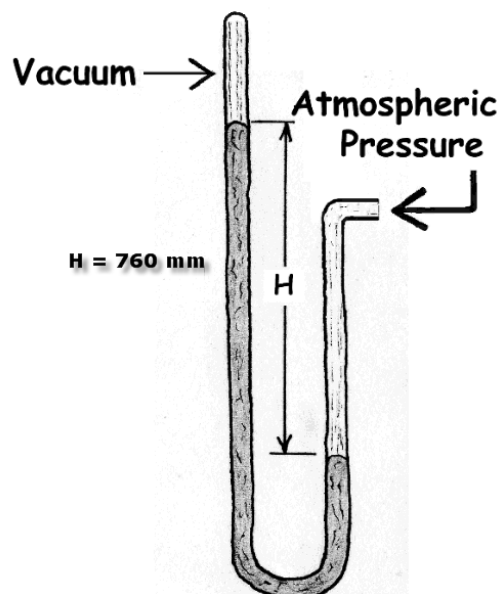


<http://wright.nasa.gov/airplane/Images/pressure.gif>

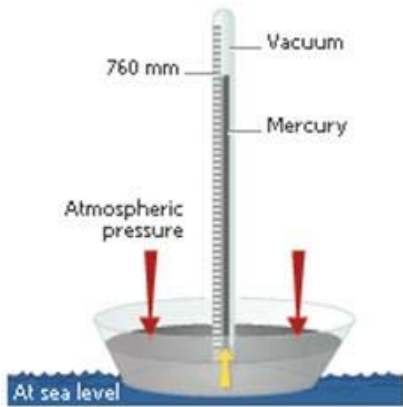
- In SI units, this pressure equals 1.013×10^5 Pa.
- The equivalence between units is:

$$1 \text{ atm} = 760 \text{ mm Hg} = 1.013 \times 10^5 \text{ Pa}$$

- The **standard atmospheric pressure** corresponds to the typical pressure at sea level.
- This pressure is sufficient to support a column of mercury 760 mm in height.



<http://www.challengers101.com/images/Pressure25.gif>



http://webserver.dmt.upm.es/~isidoro/lab1/Piezometry/Piezometry_archivos/image009.jpg

The atmospheric pressure can support a column of 760 mm of mercury.

In this experiment, the atmospheric pressure equals the pressure exerted by the weight of mercury.

The atmospheric pressure can crush a can: if we heat the water in a can, until steam is formed and then close it and let the steam to condense, then the pressure inside the can will be smaller than the external pressure and the can will crush.



http://physics.usc.edu/demolab/f1_6a.jpg



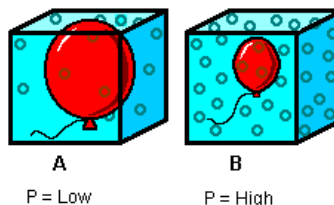
http://physics.usc.edu/demolab/f1_6b.jpg

The Pressure-Volume Relationship: Boyle's Law

If the pressure on a balloon is decreased, the balloon expands.

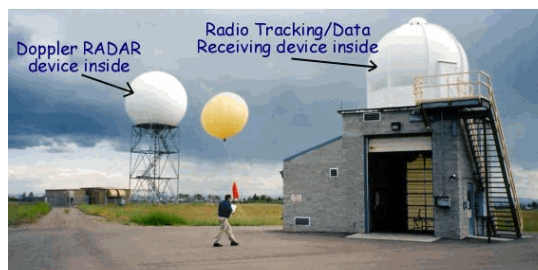
Figure 2. Volume of One Mole of Gas Under Different Conditions

All Balloons contain one mole of gas (6.02×10^{23} molecules)

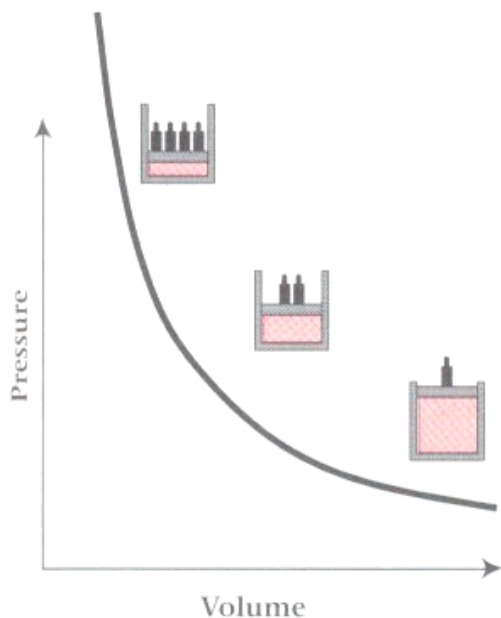


http://www.epa.gov/eogapti1/module2/idealgas/images/fig2_gas.gif

That is why weather balloons expand as they rise through the atmosphere.



http://epod.usra.edu/archive/images/main_balloon.jpg



http://www.pharmacology2000.com/physics/Chemistry_Physics/boyle1ga.gif

Boyle's law states that the volume of a fixed quantity of gas maintained at constant temperature is inversely proportional to the pressure.

Boyle's law can be expressed as

$$V = \text{constant} * \frac{1}{P} \leftrightarrow P * V = \text{constant}$$

When the pressure on a gas changes from P_1 to P_2 , the volume also changes, according to this expression (if the temperature and the number of moles remain constant):

$$[P_1 * V_1 = P_2 * V_2]_{T,n \text{ constant}}$$

Exercise

A fixed quantity of gas at 23 °C exhibits a pressure of 748 mm Hg and occupies a volume of 10.3 L. Use Boyle's law to calculate the volume the gas will occupy at 23 °C if the pressure is increased to 1.55 atm.

Sol: 6.54 L

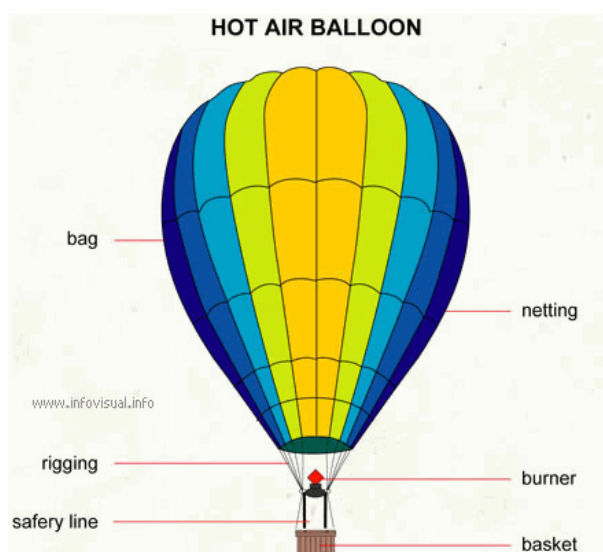
The Temperature-Volume Relationship: Charles's Law

Hot air balloons rise because air expands as it is heated. The warm air in the balloon is less dense than the surrounding cool air at the same pressure.

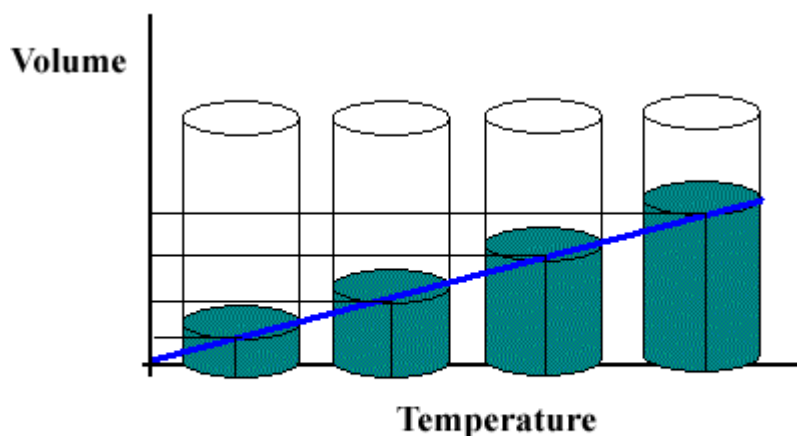
This difference in density causes the balloon to ascend. Conversely, a balloon will shrink when the gas in it is cooled.

Charles found that the volume of a fixed quantity of gas at constant pressure increases linearly with temperature.

Charles's law can be stated as follows: the volume of a fixed amount of gas maintained at constant pressure is directly proportional to its absolute temperature (temperature in Kelvin degrees)



$$V = \text{constant} * T \Leftrightarrow \frac{V}{T} = \text{constant}$$



<http://chimge.unil.ch/En/mat/images/im1mat10.2.gif>

When the temperature of a gas changes from T_1 to T_2 , the volume also changes, according to this expression:

$$\left[\frac{V_1}{T_1} = \frac{V_2}{T_2} \right]_{P,n \text{ constant}}$$

The Ideal-Gas Equation

The volume of a gas is related to the pressure (P), temperature (T) and the amount of gas (n; number of moles) in this form:

$$V \propto \frac{n \cdot T}{P}$$

If we call the proportionality constant R, we obtain:

$$V = R \left(\frac{n \cdot T}{P} \right) \rightarrow P \cdot V = n \cdot R \cdot T$$

P ... pressure ... atm

V ... volume ... L

n ... # of moles ... mol

R ... constant ... $0.082 \frac{\text{atm} \cdot \text{L}}{\text{K} \cdot \text{mol}}$

T ... temperature ... K

The volume that occupies 1 mol of a gas (regardless of the nature of that gas) when the conditions are standard temperature and pressure (STP) is:

$$V = \frac{n \cdot R \cdot T}{P} = \frac{1 \text{ mol} \cdot 0.082 \text{ atm} \cdot \text{L} / \text{K} \cdot \text{mol} \cdot 273 \text{ K}}{1 \text{ atm}} = 22.4 \text{ L}$$

Gas Mixtures and Partial Pressures

The total pressure of a mixture of gases equals the sum of the pressures that each exert if it were present alone.

The pressure exerted by a particular component of a mixture of gases is called the partial pressure of that gas.

If each of the gases obeys the ideal gas equation, we can write:

$$P_1 = n_1 \cdot \left(\frac{R \cdot T}{V} \right); P_2 = n_2 \cdot \left(\frac{R \cdot T}{V} \right); P_3 = n_3 \cdot \left(\frac{R \cdot T}{V} \right) \dots$$

We obtain that the total pressure is:

$$P_T = P_1 + P_2 + P_3 + \dots = (n_1 + n_2 + n_3 + \dots) \frac{R \cdot T}{V} = n_T \cdot \frac{R \cdot T}{V}$$

EXERCISES (PROBLEMS)

An spray can with a volume of 250 mL contains 2.3 g of propane gas ($\text{CH}_3\text{-CH}_2\text{-CH}_3$). If the can is at 18°C , what is the pressure in the can?

Atomic mass: C=12; H=1.

SOL: 4.96 atm

Calculate the number of molecules in a deep breath of air whose volume is 2.55 L at body temperature, 37°C , and a pressure of 740 mmHg.

SOL: 5.88×10^{22} gas molecules

Calculate the density of NO_2 gas at 0.97 atm and 35°C

Atomic mass: N=14; O=16

SOL: 1.77 g/L