# Gases

# Properties of Gases

**assume the volume and shape of their containers**

**most compressible of the states of matter**

**mix evenly and completely with other gases**

**much lower density than other forms of matter**

# Substances that Exist as Gases

| Group | 1 | 2 | | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| | 1A | 2A | | 3B | 4B | 5B | 6B | 7B | | 8B | | 1B | 2B | 3A | 4A | 5A | 6A | 7A | 8A |
| Period | | | | | | | | | | | | | | | | | | | |
| 1 | 1 H | | | | | | | | | | | | | | | | | | 2 He |
| 2 | 3 Li | 4 Be | | | | | | | | | | | | 5 B | 6 C | 7 N | 8 O | 9 F | 10 Ne |
| 3 | 11 Na | 12 Mg | | | | | | | | | | | | 13 Al | 14 Si | 15 P | 16 S | 17 Cl | 18 Ar |
| 4 | 19 K | 20 Ca | | 21 Sc | 22 Ti | 23 V | 24 Cr | 25 Mn | 26 Fe | 27 Co | 28 Ni | 29 Cu | 30 Zn | 31 Ga | 32 Ge | 33 As | 34 Se | 35 Br | 36 Kr |
| 5 | 37 Rb | 38 Sr | | 39 Y | 40 Zr | 41 Nb | 42 Mo | 43 Tc | 44 Ru | 45 Rh | 46 Pd | 47 Ag | 48 Cd | 49 In | 50 Sn | 51 Sb | 52 Te | 53 I | 54 Xe |
| 6 | 55 Cs | 56 Ba | * | 71 Lu | 72 Hf | 73 Ta | 74 W | 75 Re | 76 Os | 77 Ir | 78 Pt | 79 Au | 80 Hg | 81 Tl | 82 Pb | 83 Bi | 84 Po | 85 At | 86 Rn |
| 7 | 87 Fr | 88 Ra | ** | 103 Lr | 104 Rf | 105 Db | 106 Sg | 107 Bh | 108 Hs | 109 Mt | 110 Uun | 111 Uuu | 112 Uub | 113 Uut | 114 Uuq | 115 Uup | 116 Uuh | 117 Uus | 118 Uuo |
| lanthanides | | | * | 57 La | 58 Ce | 59 Pr | 60 Nd | 61 Pm | 62 Sm | 63 Eu | 64 Gd | 65 Tb | 66 Dy | 67 Ho | 68 Er | 69 Tm | 70 Yb | | |
| actinides | | | ** | 89 Ac | 90 Th | 91 Pa | 92 U | 93 Np | 94 Pu | 95 Am | 96 Cm | 97 Bk | 98 Cf | 99 Es | 100 Fm | 101 Md | 102 No | | |

**Elements that exist as gases at 25˚C and 1 atm. The Noble gases (the Group 8A elements) are monatomic species; the other elements exist as diatomic molecules. Ozone ($O_3$) is also a gas.**

# Some substances found as gases at 1 atm and 25 °C

**Elements**

$H_2$, $N_2$, $O_2$, $O_3$, $F_2$, $Cl_2$

He, Ne, Ar, Kr, Xe, Rn

**Compounds**

HF, HCl, HBr, HI, CO, $CO_2$, $NH_3$, NO,

$NO_2$, $N_2O$, $SO_2$, $H_2S$, HCN

# Pressure of a Gas

# Pressure

**The force exerted on an object divided by the surface area of the object;**

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

*Any gas confined to a container is found to exert a pressure on the container .*

# SI Units of Pressure

**customary units**

**1 standard atm = 760 mm Hg**

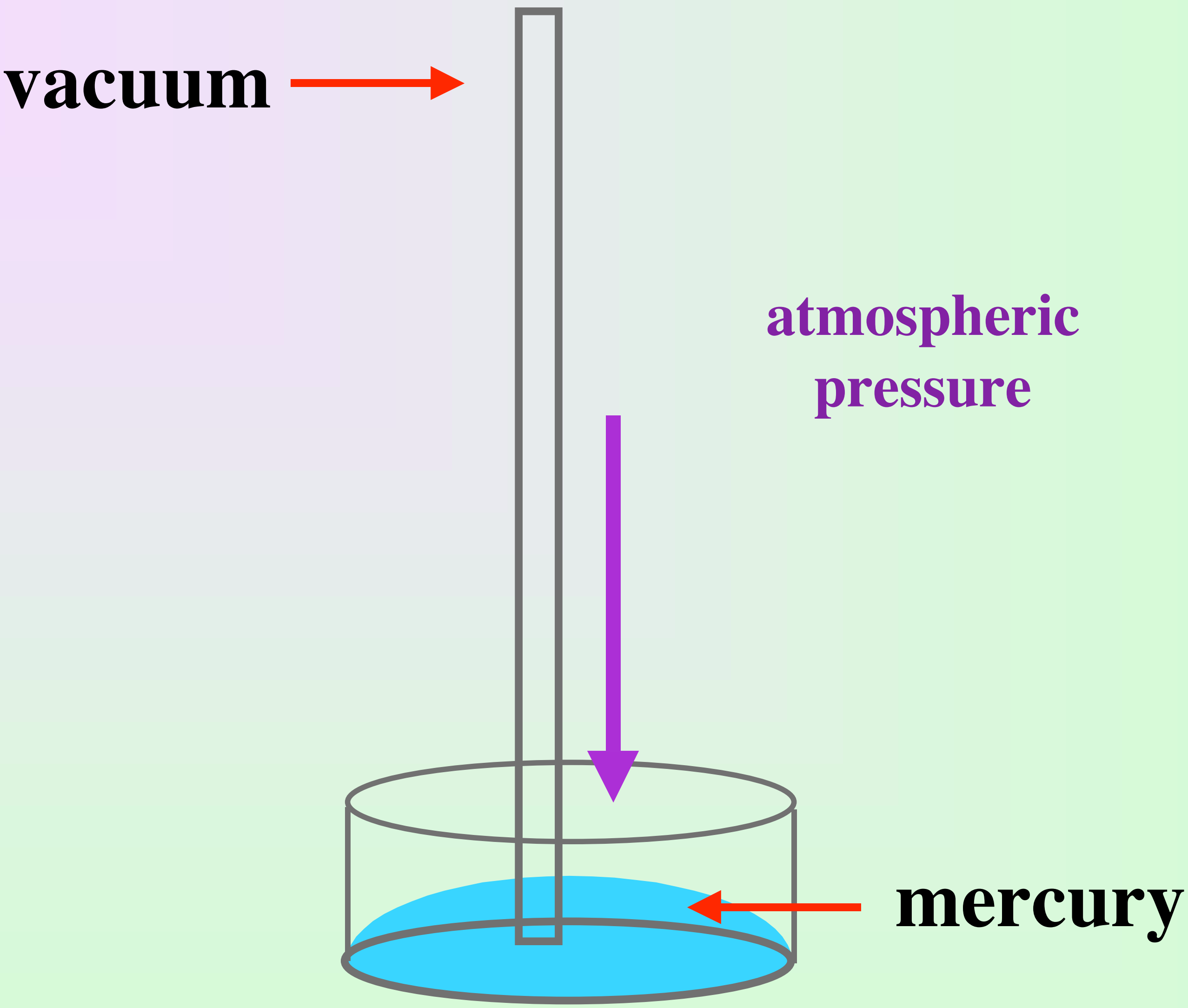
**= 760 torr**

**SI units**

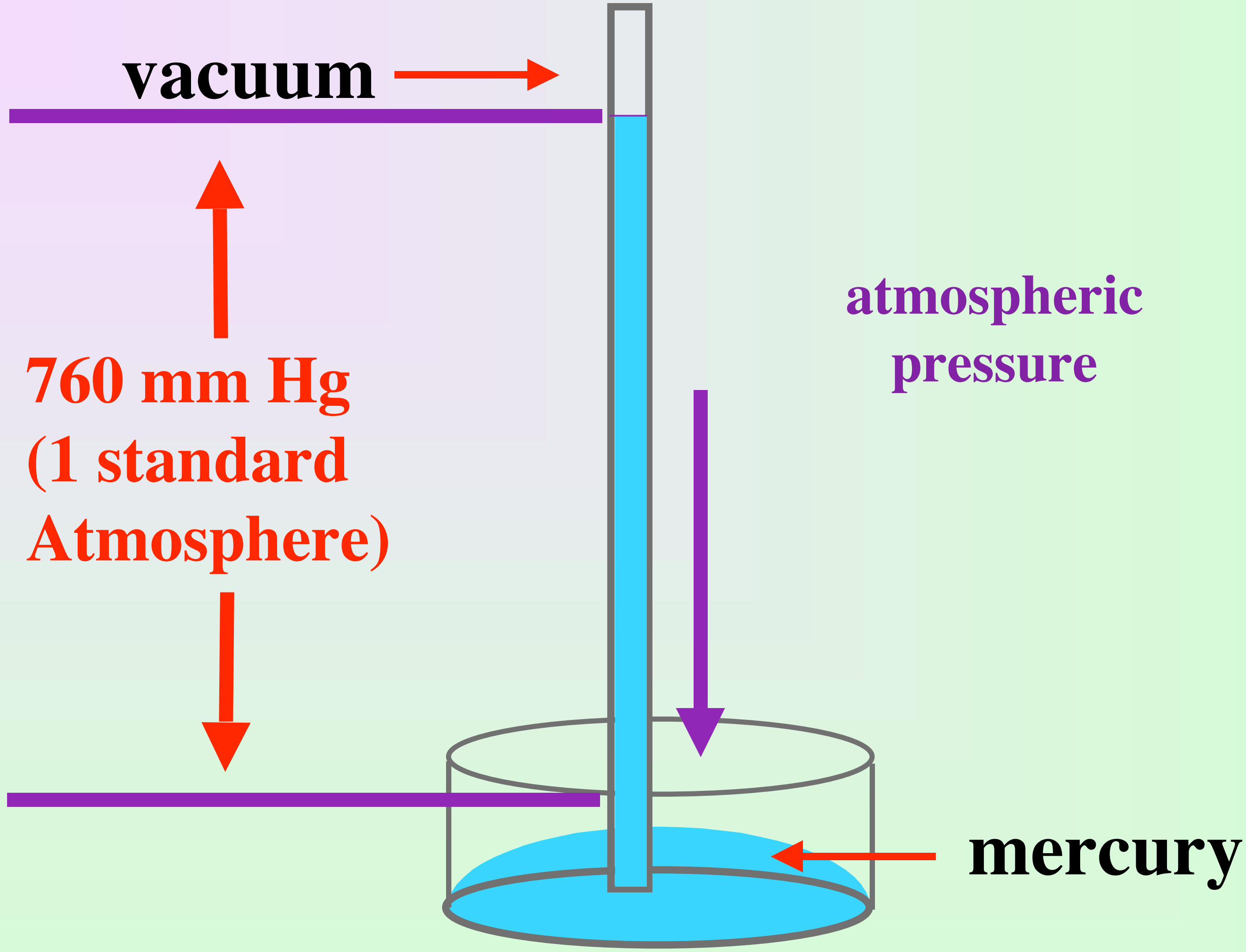
**pressure = force/area**

**pressure = Newton/$m^2$ = Pascal**

**1 standard atmosphere = 101,325 Pa**




A Torricellian Barometer




A Torricellian Barometer

# The Gas Laws

**Boyle's Law**

**Charles' Law**

**Avogadro's Law**

**The Ideal Gas Law**

# The Gas Laws and the Scientific Method

**Observations**

**Laws** — **Boyle, Charles, Avogadro, Ideal Gas**

**Theory** — **Kinetic-molecular theory**

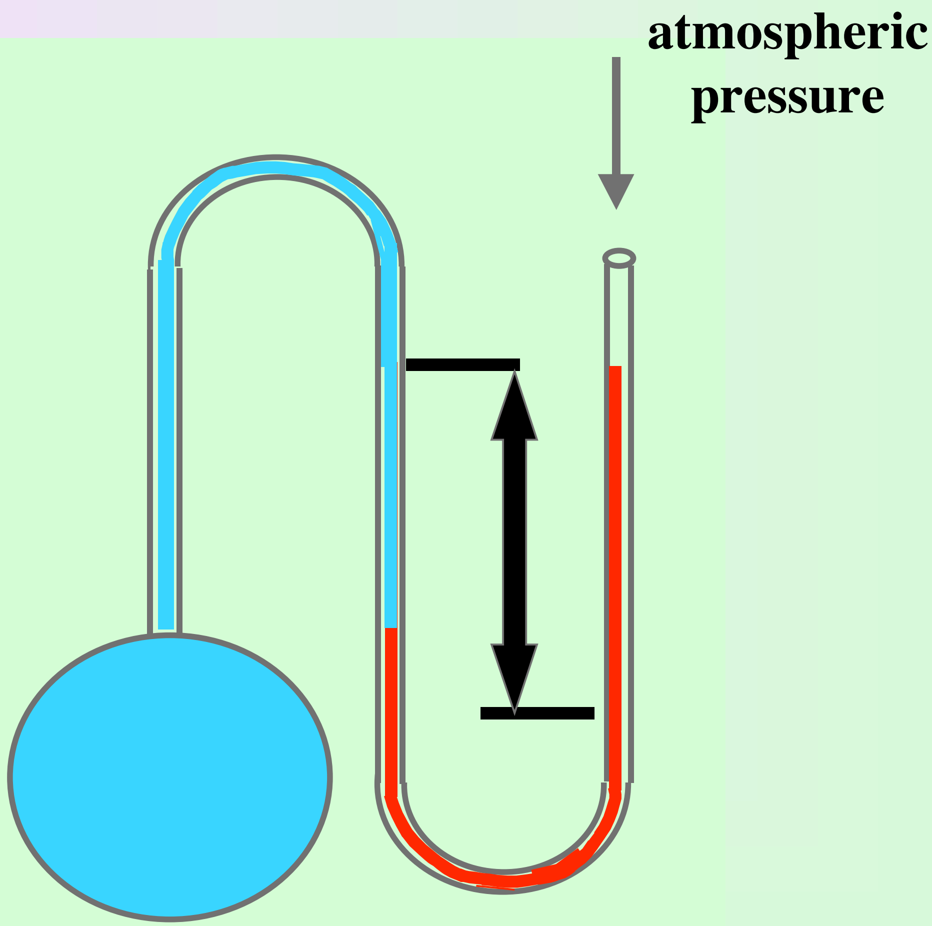
**Hypothesis**

**Experiment**

# Boyle's Law

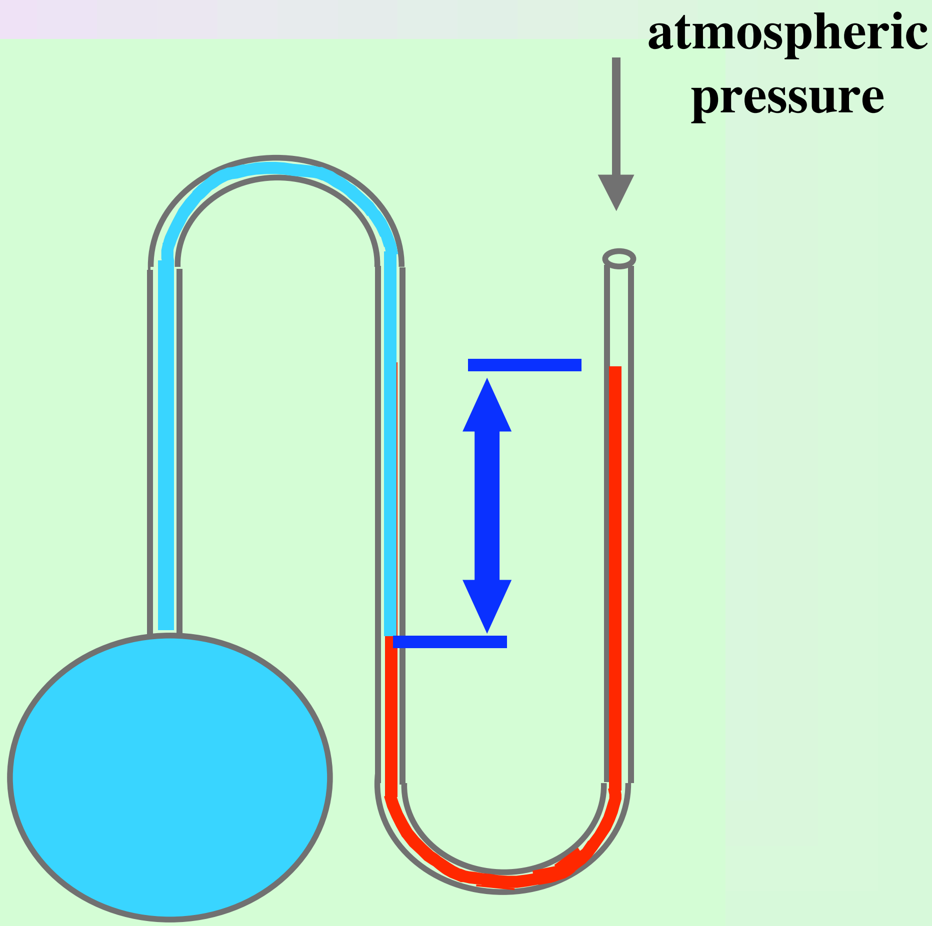
***Robert Boyle 1626 – 1691***

# Manometer




$$P_{gas} = P_{atm} - h_{Hg}$$

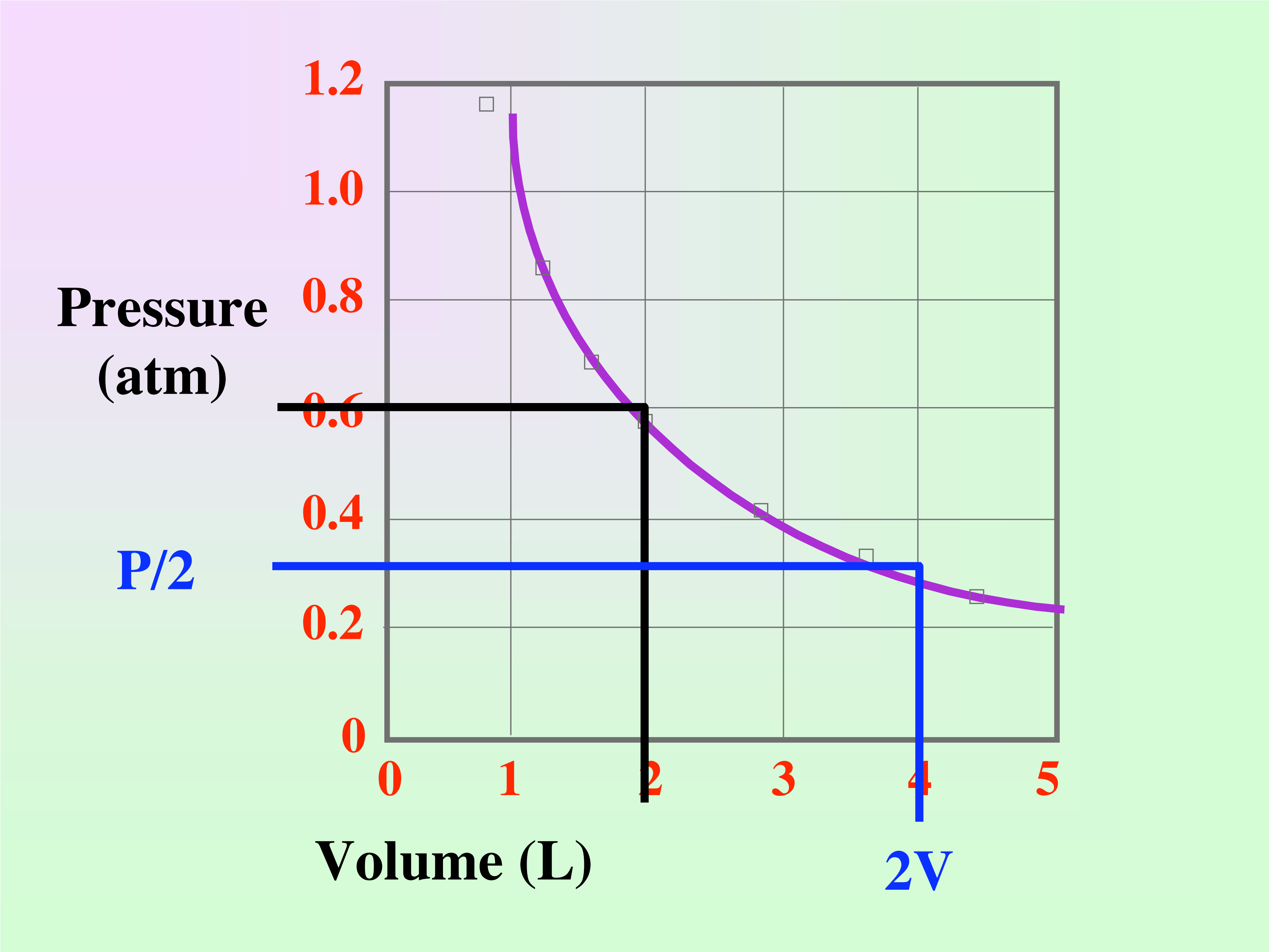
# Manometer




$$P_{gas} = P_{atm} + h_{Hg}$$

# Boyle's Data

| Pressure (mm Hg) | Volume (arbitrary units) |
|---|---|
| 724 | 1.50 |
| 869 | 1.33 |
| 951 | 1.22 |
| 998 | 1.16 |
| 1230 | 0.94 |
| 1893 | 0.61 |
| 2250 | 0.51 |



Pressure
(atm)
1.2
1.0
0.8
0.6
0.4
P/2
0.2
0
0
1
2
3
4
5
Volume (L)
2V

# Boyle's Data

| Pressure (mm Hg) | Volume (arbitrary units) | PV |
|---|---|---|
| 724 | 1.50 | 1090 |
| 869 | 1.33 | 1160 |
| 951 | 1.22 | 1160 |
| 998 | 1.16 | 1160 |
| 1230 | 0.94 | 1200 |
| 1893 | 0.61 | 1200 |
| 2250 | 0.51 | 1100 |

# Boyle's Law

**at constant temperature, the volume of a constant amount of gas is inversely proportional to the pressure**

**constant *n*, constant *T***

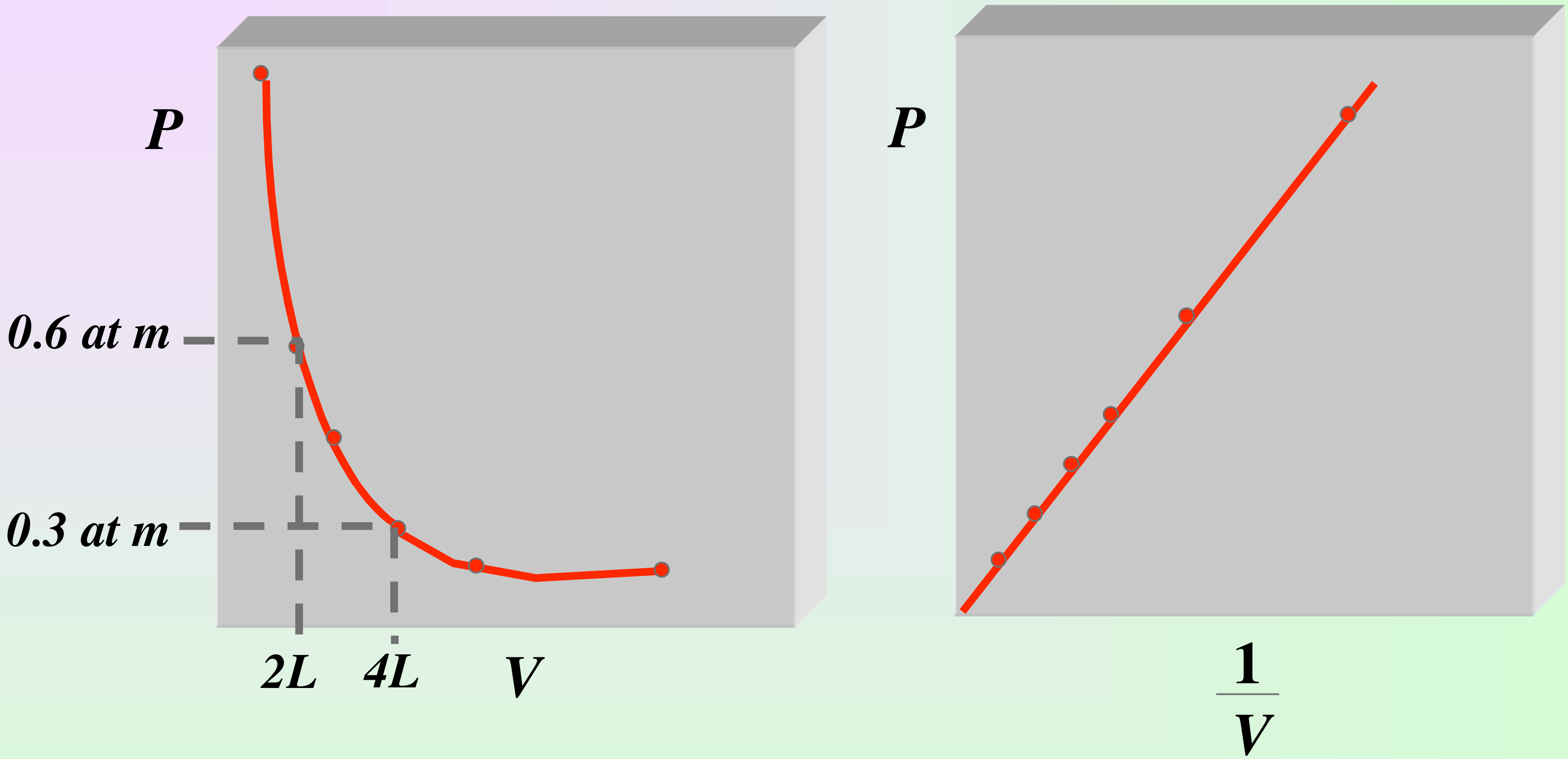
$$PV = k$$

# Plot of *P* versus *V* is a hyperbola

$PV = k$

$P = (1/V)\ k$

**is of the form y = mx + b, which is the equation for a straight line**



(a) (b)

# Charles's Law

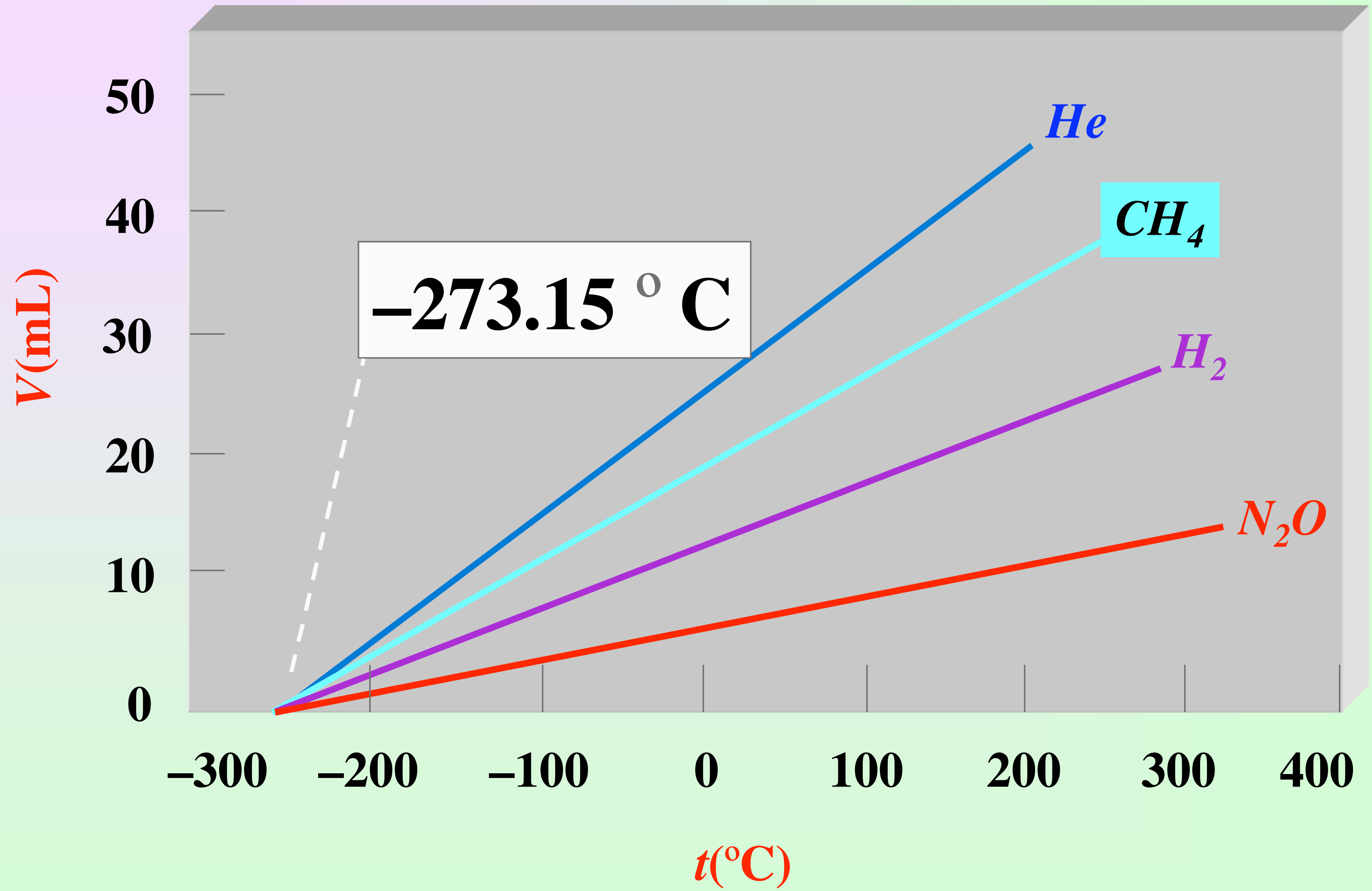
*Jacques Charles 1746 – 1823*

# Charles's Law

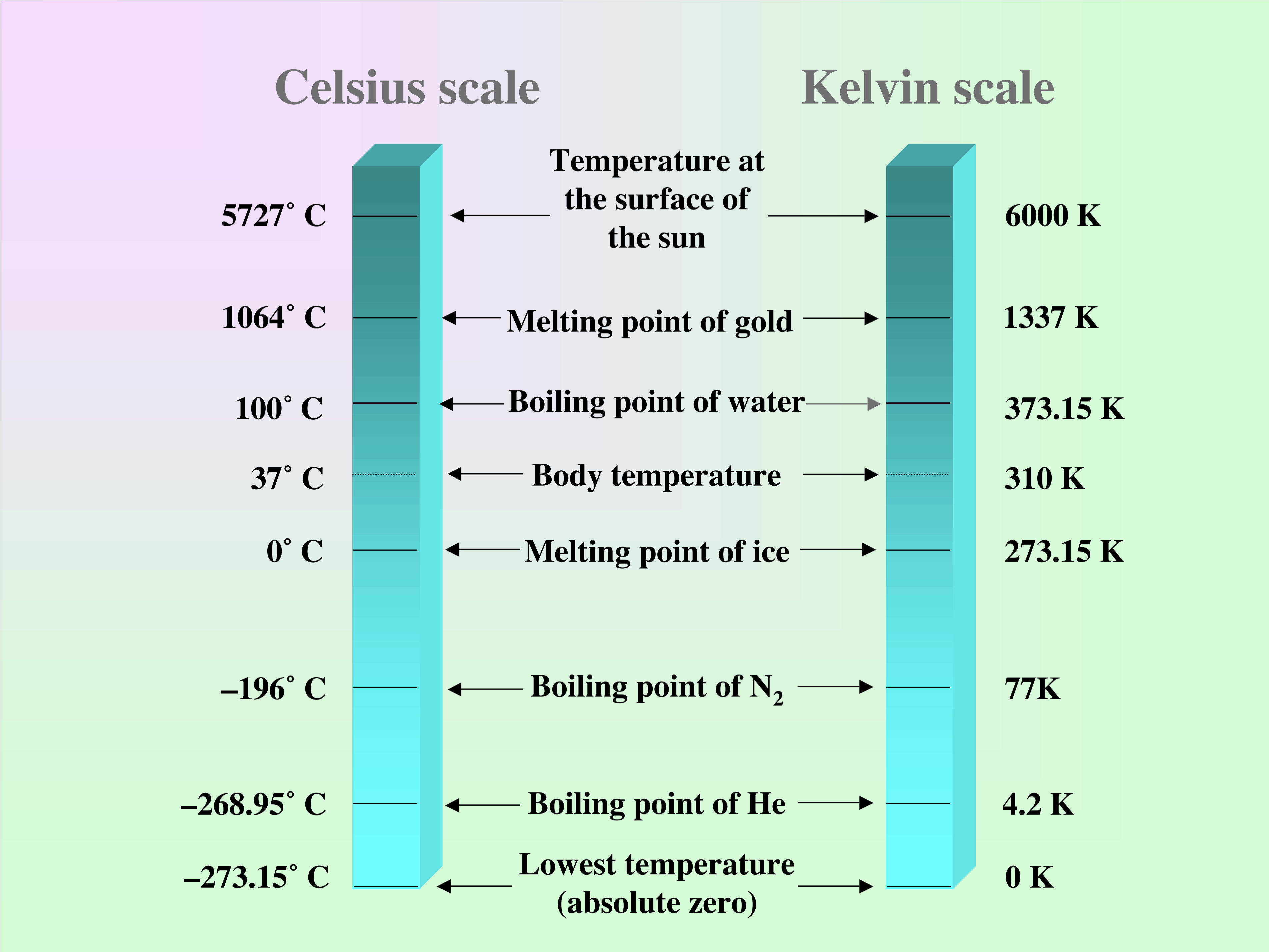
**at constant pressure, the volume of a constant amount of gas is directly proportional to the absolute temperature**

**constant *n*, constant *P***

$$V = kT$$



V(mL)
50
40
30
20
10
0
–273.15 °C
He
CH4
H2
N2O
–300
–200
–100
0
100
200
300
400
t(°C)



Celsius scale
Kelvin scale
5727° C
Temperature at the surface of the sun
6000 K
1064° C
Melting point of gold
1337 K
100° C
Boiling point of water
373.15 K
37° C
Body temperature
310 K
0° C
Melting point of ice
273.15 K
–196° C
Boiling point of N2
77K
–268.95° C
Boiling point of He
4.2 K
–273.15° C
Lowest temperature (absolute zero)
0 K

# Avogadro's Law

***Amadeo Avogadro (1776 – 1856)***

# Avogadro's Law

**at constant temperature and pressure, the volume of a gas is directly proportional to the number moles**

**constant T, constant P**

$$V = kn$$

**equal volumes of different gases contain equal numbers of molecules**

# The Ideal-Gas Equation

# The Ideal Gas Law

$$PV = n\mathrm{R}T$$

**constant *n*, constant *T* (Boyle's Law)**

**constant *n*, constant *P* (Charles' Law)**

**constant *P*, constant *T* (Avogadro's Law)**

# The Ideal-Gas Law

$PV = nRT$

**P is pressure in atmospheres**
**V is volume in liters**
**n is number of moles**
**T is absolute temperature in Kelvins**
**R is called the gas constant**

**An ideal gas is a hypothetical gas whose pressure-volume-temperature behavior can be completely accounted for by the ideal-gas equation.**

# Standard temperature and pressure

$PV = nRT$

**standard conditions defined**

| | |
|---|---|
| $P$ = 1 atm<br>$T$ = 0°C (273 K) | *STP* |

# Standard Molar Volume

**The volume occupied by 1 mole of a gas at STP is 22.414 L**

# The Gas Constant R

**The volume occupied by 1 mole of a gas at STP is 22.414 L**

$$PV = nRT$$

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

$$R = \frac{(1\ \text{atm})\ (22.414\ \text{L})}{(1\ \text{mol})\ (273.15\ \text{K})}$$

$$R = 0.082057\ (\text{L atm/mol K})$$

# Example

**What is the volume occupied by 49.8 g of HCl at STP?**

$$49.8\ \text{g} \times \frac{1\text{mol}}{36.46\ \text{g}} = 1.366\ \text{mol}$$

$P$ = 1 atm

$T$ = 273 K

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

$$\frac{(1.366\ \text{mol})\ (0.0821\ \text{L atm/mol K})\ (273\text{K})}{1\ \text{atm}}$$

V = 30.6 L

# Example

**What is the volume occupied by 49.8 g of HCl at STP? (alternative solution)**

$P = 1$ atm

$T = 273$ K

$$n = \frac{49.8\ g}{36.46\ g/mol} = 1.366\ mol$$

$$V = \frac{22.4\ \text{L}}{1\ \text{mol}} \times 1.366\ \text{mol}$$

$$V = 30.6\ \text{L}$$

# Example

**A compound has the empirical formula $BH_3$. At 27 °C, 74.3 mL of the gas exerted a pressure of 1.12 atm. If the mass of the gas was 0.0934 g, what is its molecular formula?**

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

# Example (cont.)

**A compound has the empirical formula $BH_3$. At 27 °C, 74.3 mL of the gas exerted a pressure of 1.12 atm. If the mass of the gas was 0.0934 g, what is its molecular formula?**

$n = 0.00338$ mol

$$\frac{0.0934\ \text{g}}{0.00338\ \text{mol}} = 27.6\ \text{g/mol}$$

$BH_3 = 13.8$ g / empirical formula

$B_2H_6$

## The ideal gas law is often used to calculate the changes that will occur when the conditions of a gas are changed

$$PV = nRT$$

*If nRT are constant*

$$P = (nRT)\frac{1}{V}$$

$$PV = (nRT) = P_2V_2$$

$$PV = nRT$$

*If nRP are constant*

$$V = \left(\frac{nR}{P}\right)T$$

$$\frac{V}{T} = \left(\frac{nR}{P}\right) = \frac{V_2}{T_2}$$

## The ideal gas law is often used to calculate the changes that will occur when the conditions of a gas are changed

$$PV = nRT$$

*If PRT are constant*

$$V = \left(\frac{RT}{P}\right)n$$

$$\frac{V}{n} = \left(\frac{RT}{P}\right) = \frac{V_2}{n_2}$$

$$PV = nRT$$

*If nRVare constant*

$$P = \left(\frac{nR}{V}\right)T$$

$$\frac{P}{T} = \left(\frac{nR}{V}\right) = \frac{P_2}{T_2}$$

# Example

**A sample of oxygen gas initially at 0.97 atm is cooled from 21 $^{o}$ C to -68 $^{o}$ C at constant volume. What is its final pressure.**

**PV = nRT**

$$\frac{P}{T} = \left(\frac{nR}{V}\right) = \frac{P_2}{T_2}$$

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

# Example

A sample of oxygen gas initially at 0.97 atm is cooled from 21 °C to -68 °C at constant volume. What is its final pressure.

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

$$\frac{(0.97\ \text{atm})}{292\ \text{K}} = \frac{P_2}{205\ \text{K}}$$

$$P_2 = 0.68\ \text{atm}$$

# Density Calculations

# Density Calculations

$$\text{density} = \frac{\text{mass}}{V}$$

$$n = \frac{\text{mass}}{\text{Molar mass}} = \frac{g}{g/mol}$$

$$\frac{n}{V} = \frac{P}{RT}$$

$$\frac{\text{mass}}{\text{Molar mass}(V)} = \frac{P}{RT}$$

Therefore: $\text{density} = \frac{P}{RT}(\text{Molar mass})$

# Example

**What is the density of $UF_6$ gas at 62 °C and 779mmHg ?**

$$\frac{n}{V} = \frac{P}{RT}$$

$$\frac{n}{V} = \frac{\dfrac{779\text{mmHg}}{760\text{mmHg/atm}}}{(0.0821\text{ L atm/mol K})\ (335\text{ K})}$$

$$= \frac{0.0373\ \text{mol}}{\text{L}} \times \frac{352.03\text{g}}{1\ \text{mol}} = 13.1\text{g/L}$$

A common unit for gasses

# Example

**Cyanogen ,empirical formula CN, is a gas with a density of 2.335 g/L at 0°C and 1 atm. What is its molecular formula?**

$$n = \frac{P\,V}{RT}$$

$$n = \frac{(1atm)\ (1\ L)}{(0.0821\ \ L\ atm/mol\ K)\ (273K)}$$

$$= \frac{2.335g}{0.0446\ \ mol} = \frac{52\ g/mol\ Cyanogen}{26g\ /mol\ (CN)} = 2$$

$C_2N_2$