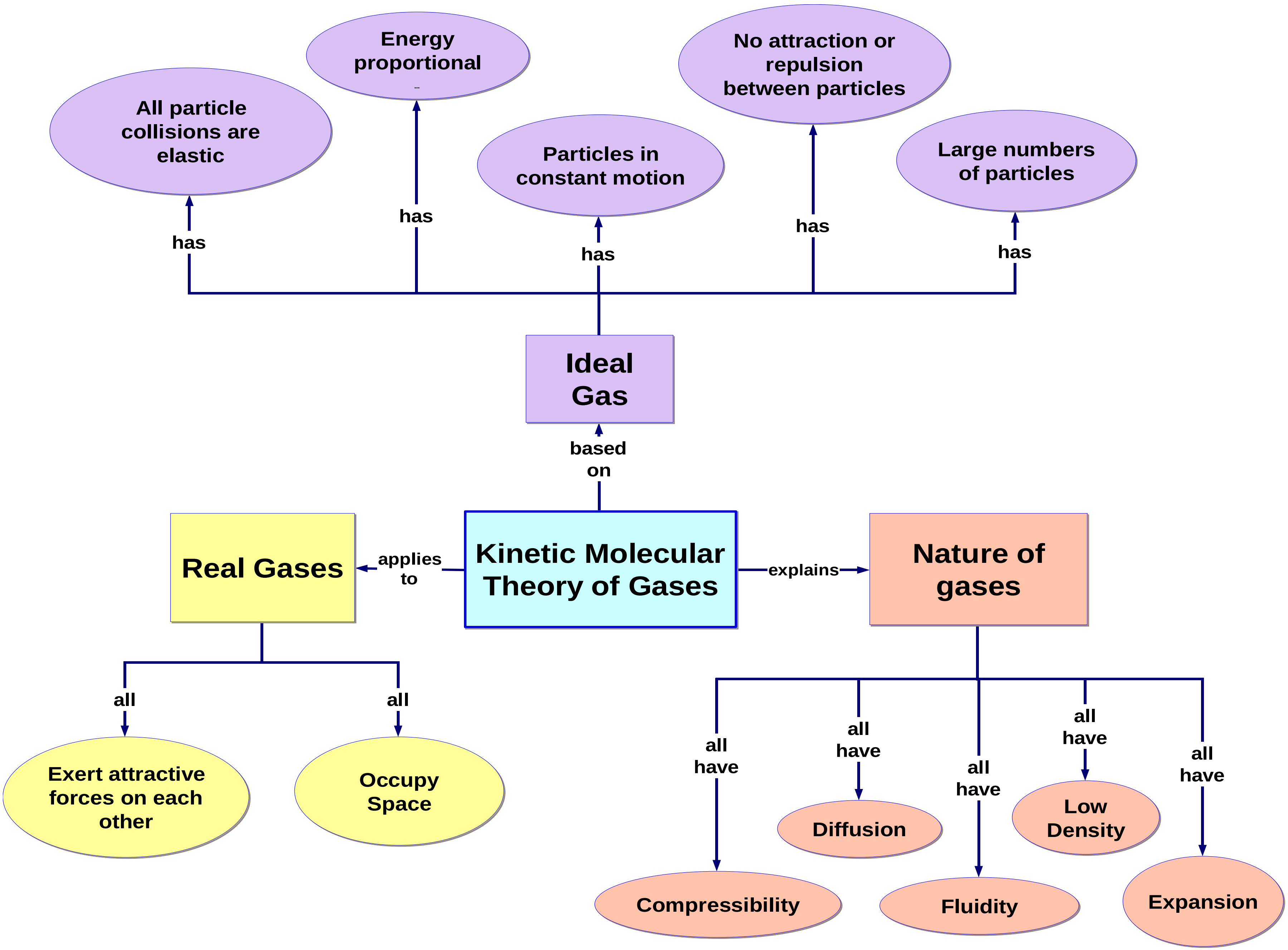
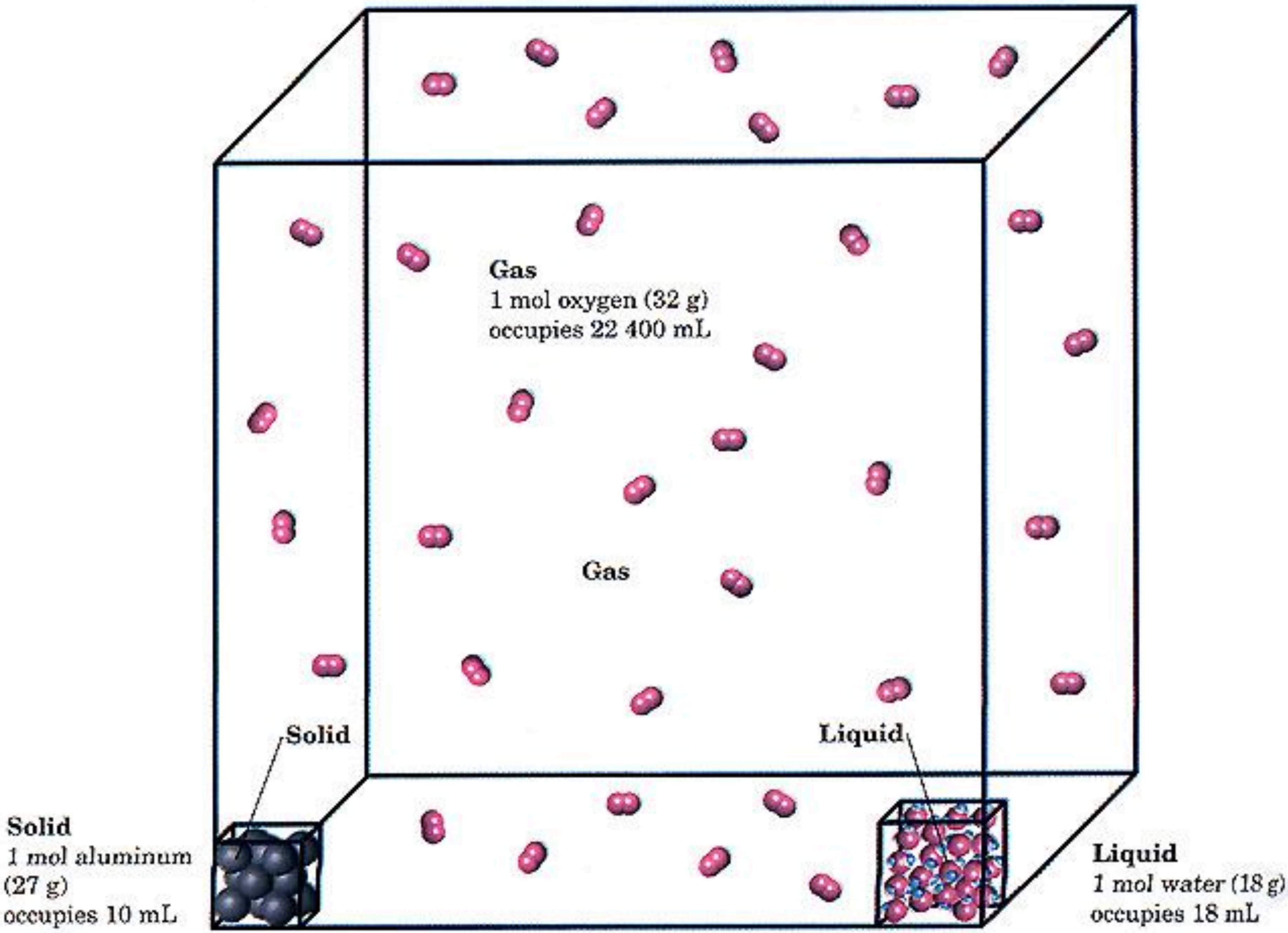
CHAPTER 12:

Behavior of Gases



All particle collisions are elastic
Energy proportional
Particles in constant motion
No attraction or repulsion between particles
Large numbers of particles
has
has
has
has
has
Ideal Gas
based on
Real Gases
applies to
Kinetic Molecular Theory of Gases
explains
Nature of gases
all
all
Exert attractive forces on each other
Occupy Space
all have
all have
all have
all have
all have
Diffusion
Low Density
Compressibility
Fluidity
Expansion


Gas
1 mol oxygen (32 g)
occupies 22 400 mL
Gas
Solid
Liquid
Solid
1 mol aluminum
(27 g)
occupies 10 mL
Liquid
1 mol water (18 g)
occupies 18 mL

Kinetic Molecular Theory

For an Ideal Gas:

- The motion of the particles is
  - Random
  - Linear
  - Elastic collisions

Real Gases

- Exert attractive forces on one another
- Occupy space

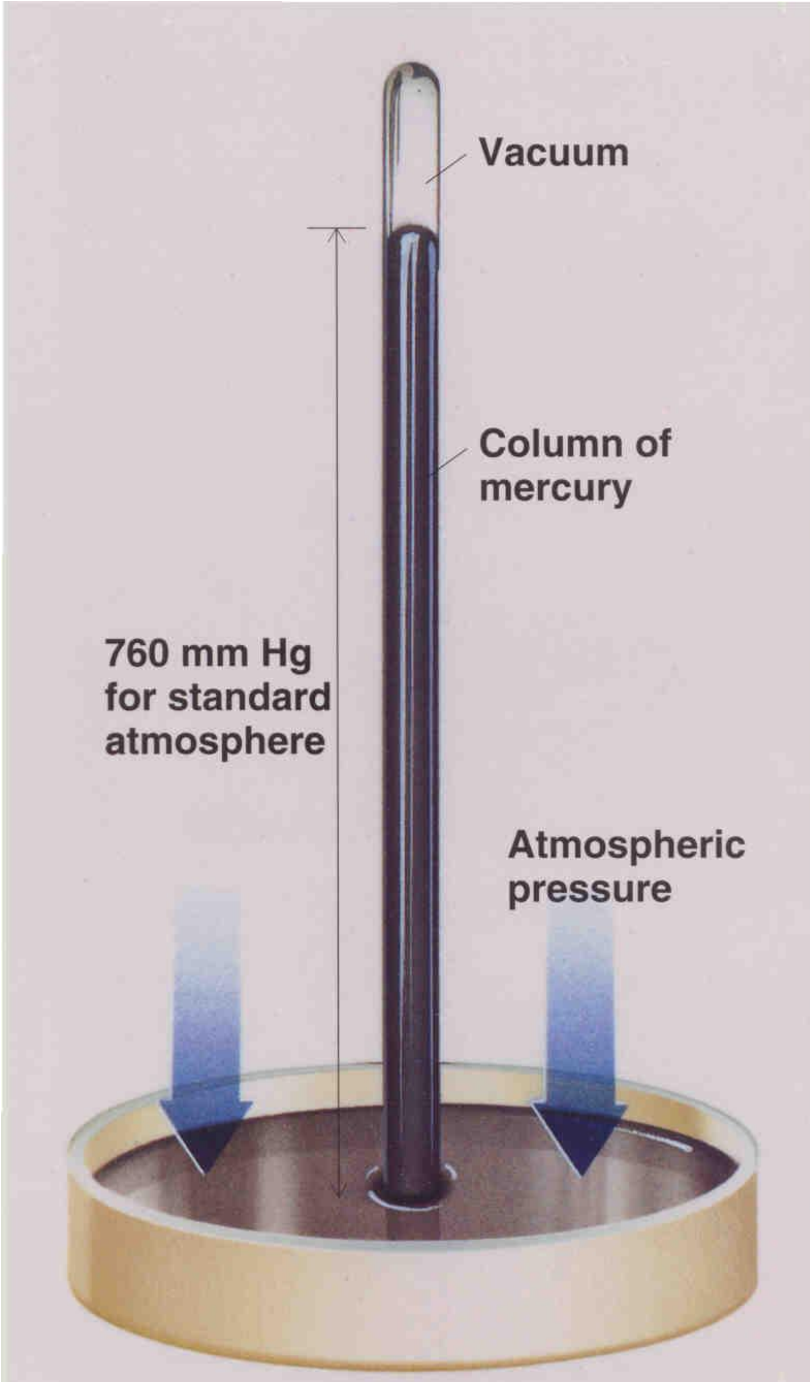
How do we define a gas?

- Volume (in L or mL)
- Pressure (atm, kPa, mm)
- Temperature (K)
- Number of Molecules (moles)

Pressure of a Gas:

- Caused by the elastic collisions in a container
- Is evenly distributed due to the randomness of particle motion

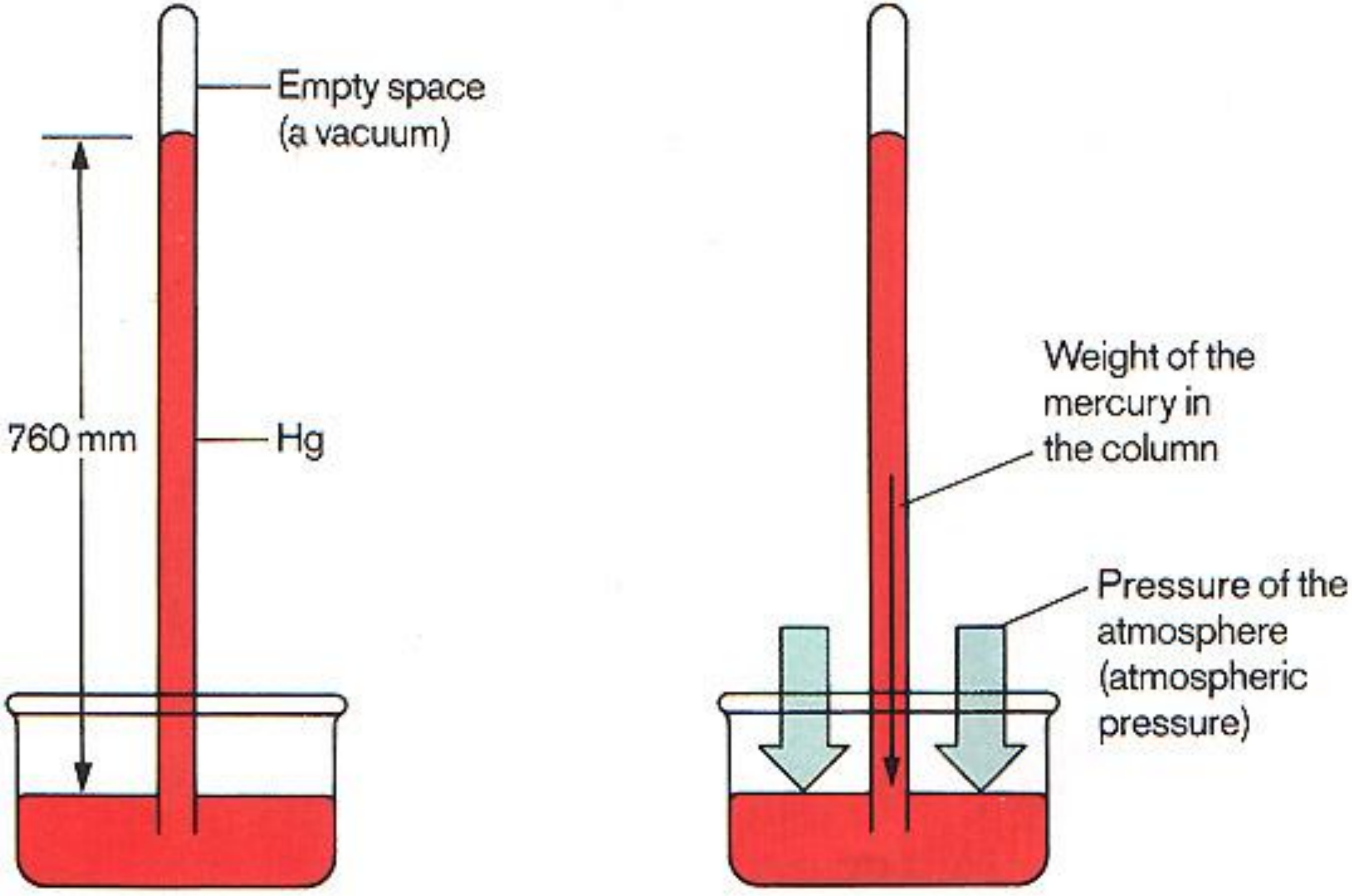
$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$



Barometer

- Device used to measure pressure
- Developed by Evangelista Torricelli

Measuring Pressure



Units of Pressure

1 Atmosphere	1 atm
1 mm Hg (torr)	1 atm = 760 mm Hg
Pascal (Pa)	1 atm = 101325 Pa
kilopascal (kPa)	1 atm = 101.325 kPa

Pressure Conversions

Review of “Pressure Conversions”

Pressure Conversions

Pressure Conversions

Conversion Factors: **1 atm = 760 mm Hg = 760 torr = 101.3 kPa**

1. How many atmospheres are 1500 mm Hg?

2. How many torr are 303.9 kPa?

3. How many mm Hg are 3 atm?

Pressure Conversions

1. How many atmospheres are 1500 mm Hg?

$$1500\ \cancel{mm}\ \text{Hg} \times \frac{1\ atm}{760\ \cancel{mm}\ \text{Hg}} = 1.973684211\ atm$$

$$= 2.0\ atm$$

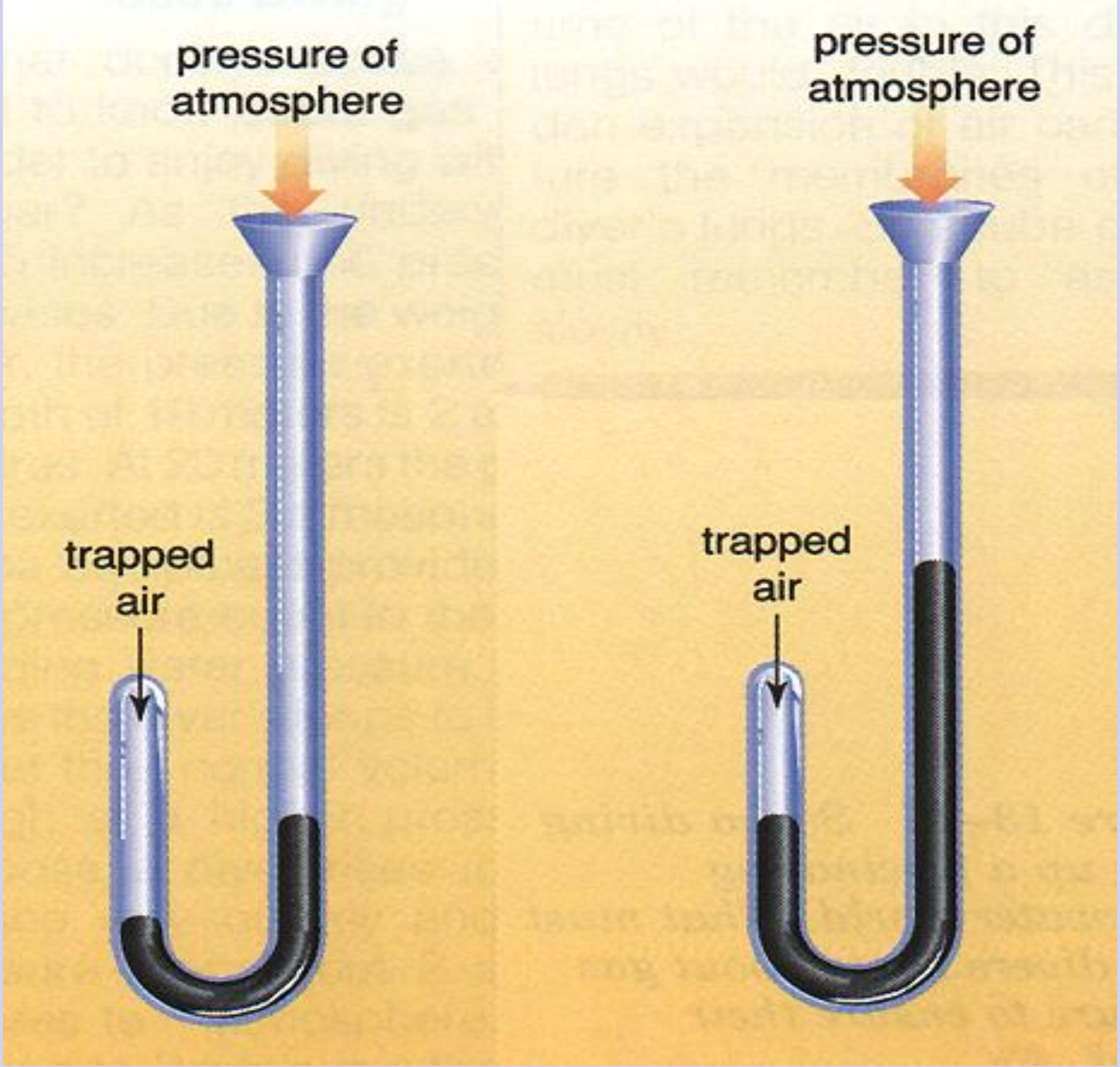
Boyle's Law

Boyle's Law

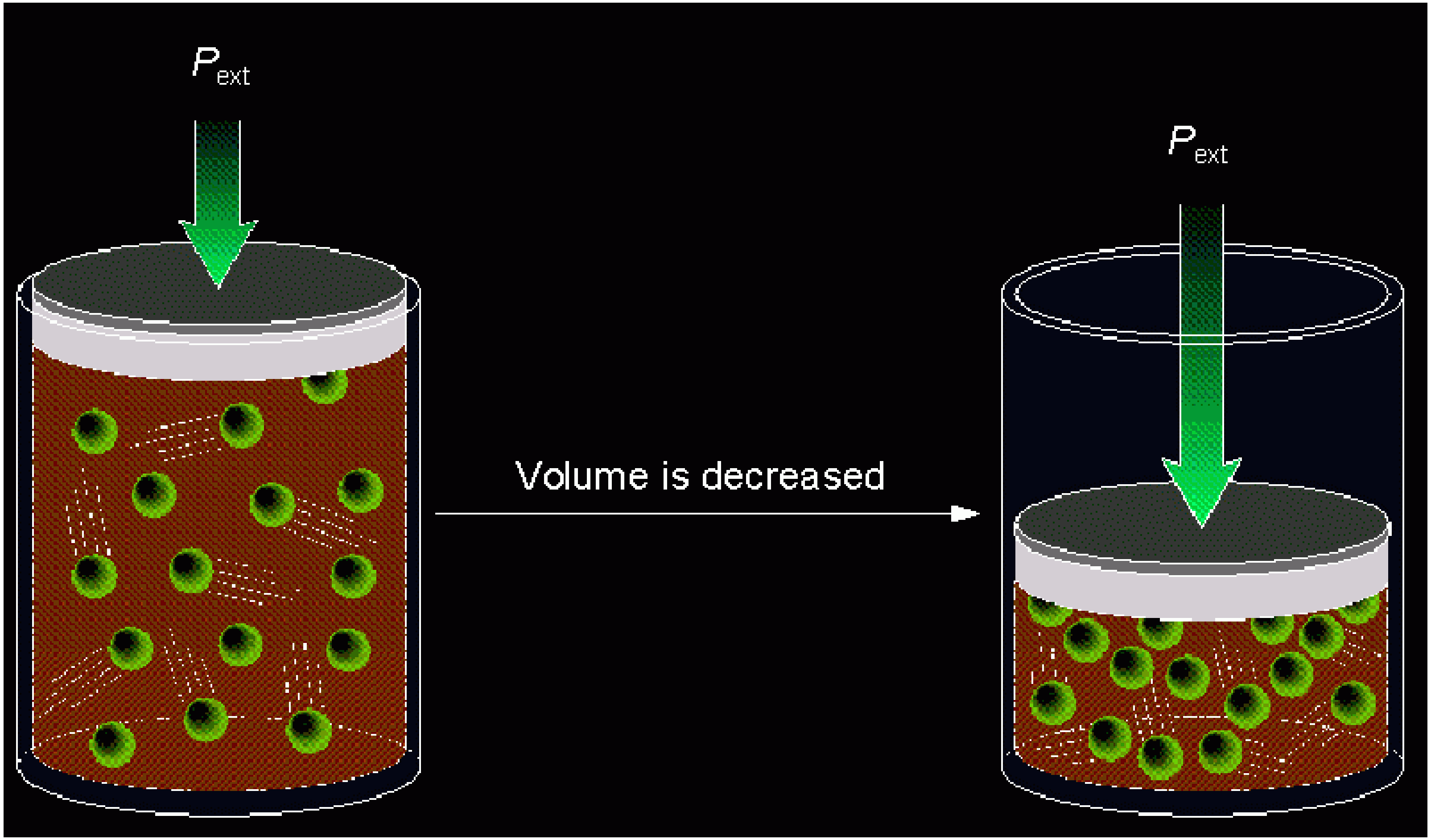
Robert Boyle experimented with the relationship between pressure and volume of gases.

Boyle's Law

He set-up a J-shaped tube and added mercury to see what it did to the volume of a trapped gas. He found that as pressure increases volume decreases.


pressure of
atmosphere
trapped
air
pressure of
atmosphere
trapped
air

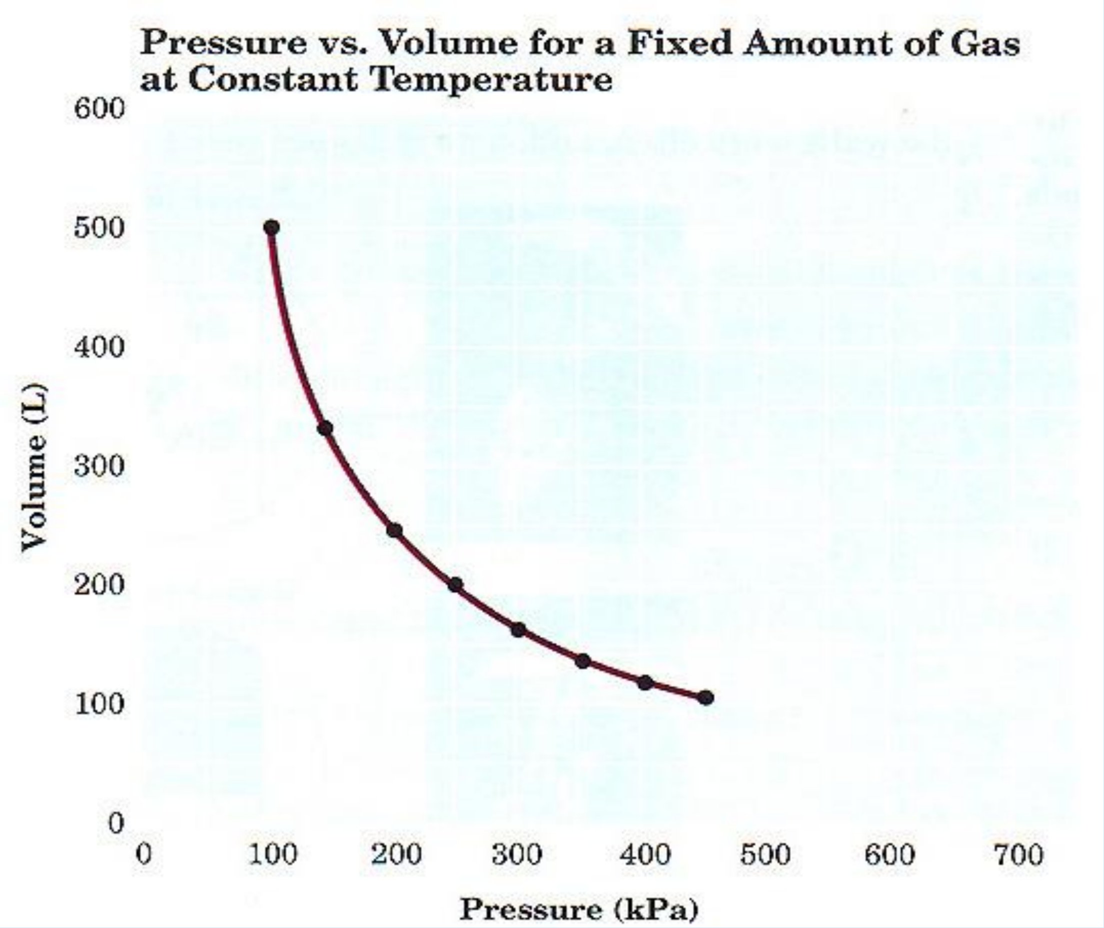
Boyle's Law



Boyles' Experiments:

Table 10-5
Pressure–Volume Data

Pressure (kPa)	Volume (mL)	*PV*
100	500	50 000
150	333	49 950
200	250	50 000
250	200	50 000
300	166	49 800
350	143	50 500
400	125	50 000
450	110	49 500



This relationship is inversely proportional, when one increases the other decreases.

Boyle's Law: PV = k

P = pressure   V = volume   k = a constant

At Constant Temperature

$$P_1V_1 = k \text{ and } P_2V_2 = k$$

$$\mathbf{P_1V_1 = k = P_2V_2}$$

So:

$$P_1V_1 = P_2V_2$$

Charles's Law Demonstration

Charles' Law

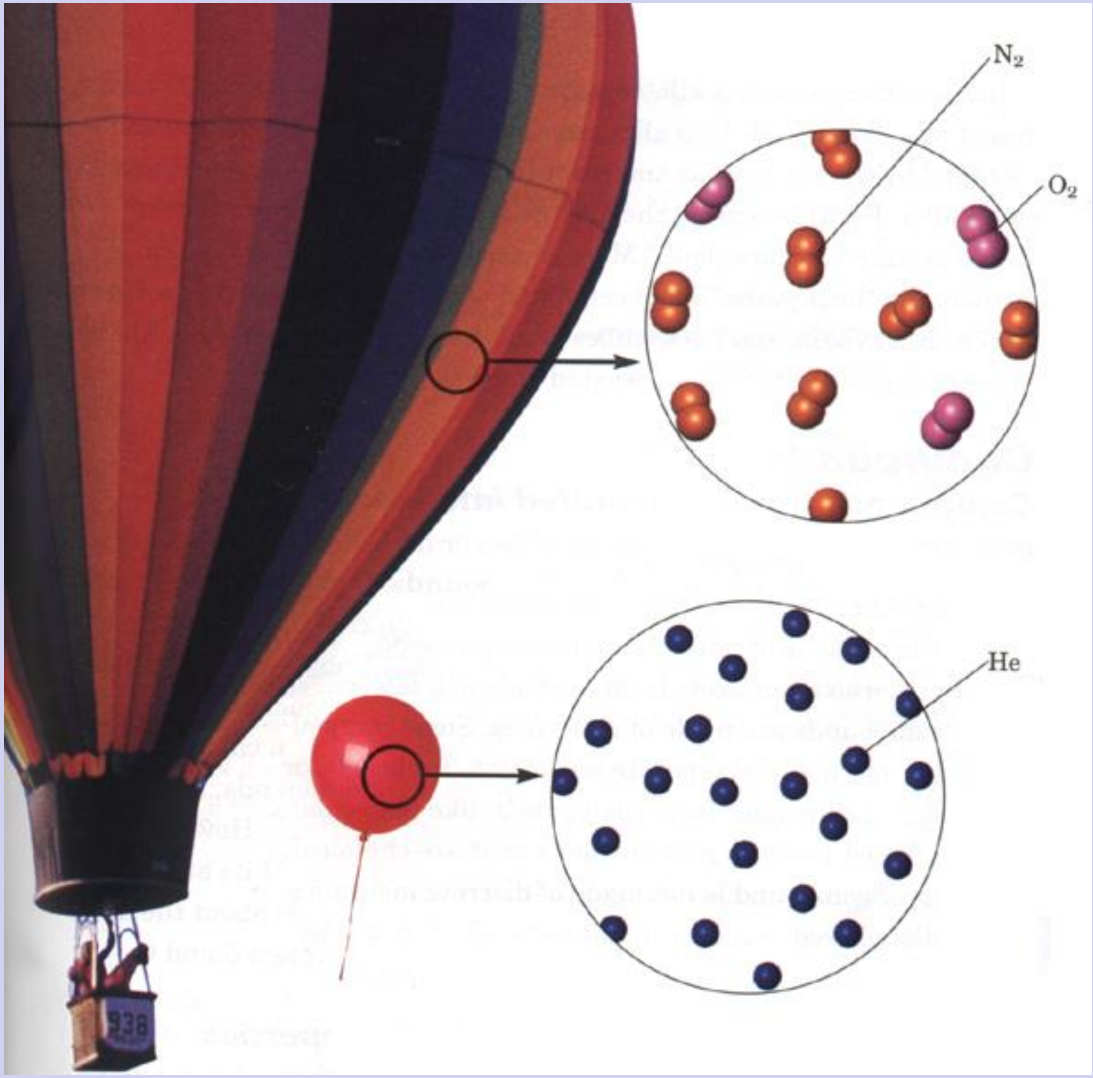
French physicist Jacques Charles was the first to fill a balloon with hydrogen gas and make a solo flight.

Charles' Law

He showed that the volume of a gas increases when the temperature increases

(at a constant pressure)

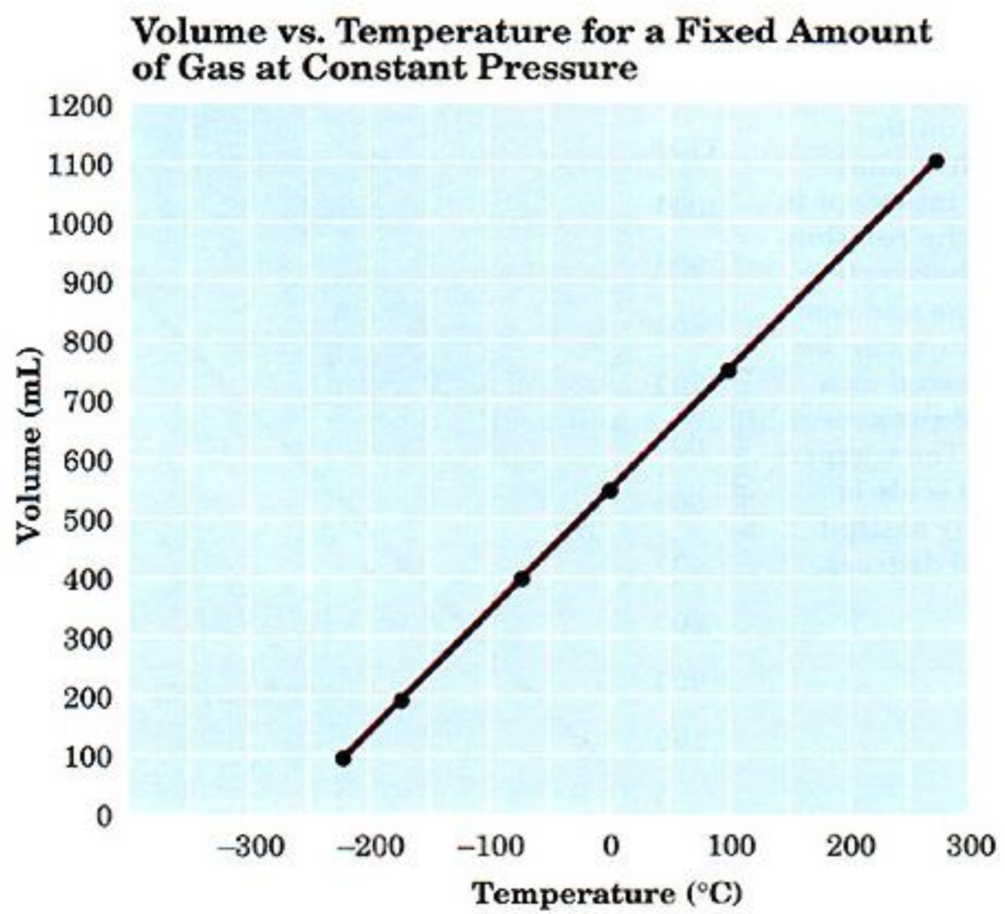
Charles' Law: Volume and Temperature



Charles' Law

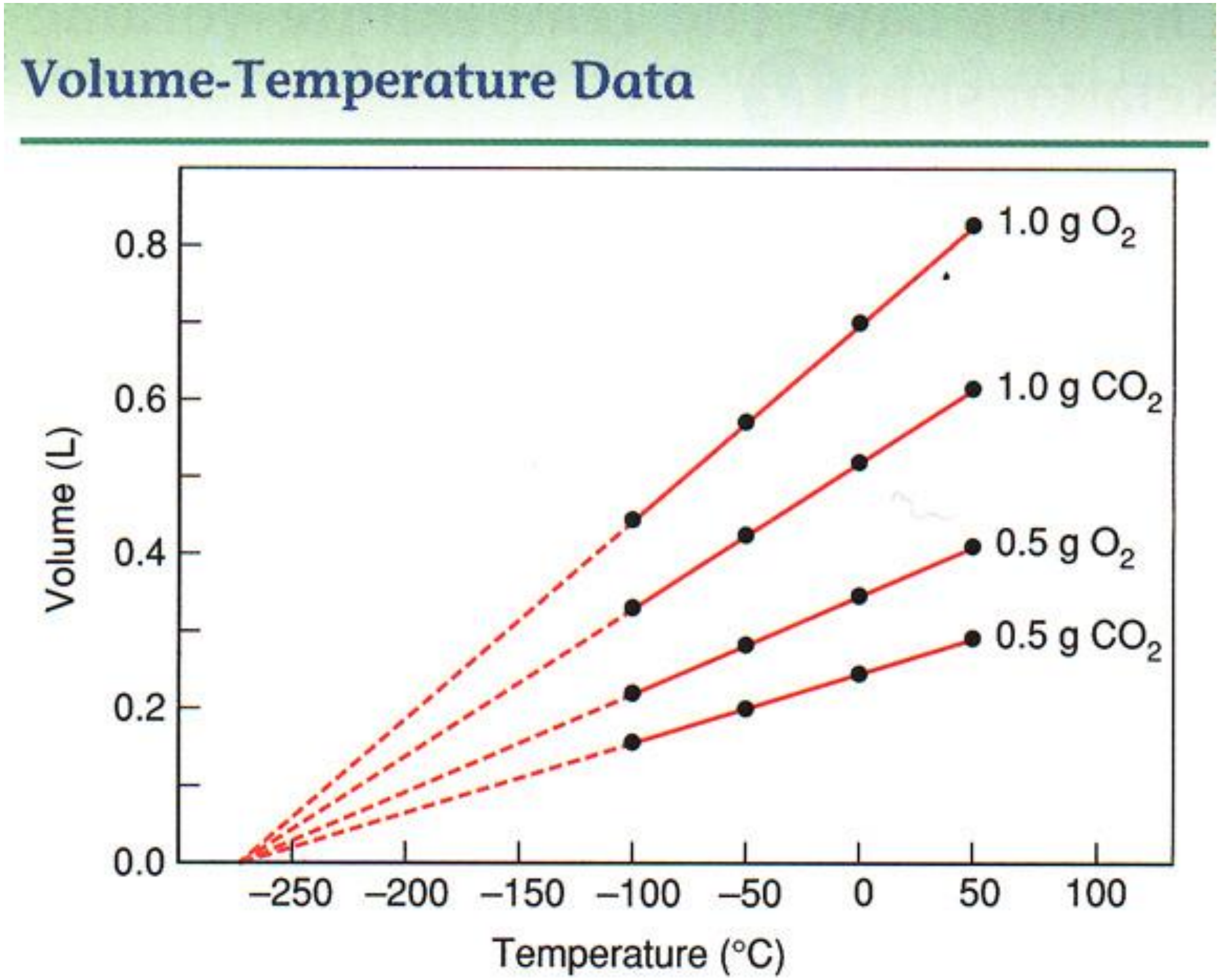
Volume–Temperature Data for a Gas at Constant Pressure

Temperature (°C)	Volume (mL)
273	1094
100	747
10	568
1	545
0	545
–1	546
–73	403
–173	199
–223	100



As temperature drops, the volume decreases. This is a linear relationship.

Volume vs. Temperature



Charles' Law:

The volume of a gas is directly proportional to the temperature

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

So: $$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$ Or: $$V_1T_2 = V_2T_1$$

Temperature must be converted to Kelvin!

Gay-Lussac’s Law

Joseph Gay-Lussac discovered that the pressure of a gas is directly proportional to the Kelvin Temperature if the volume remains constant

Gay-Lussac's Law

The pressure of a gas is directly proportional to the temperature at constant volume

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

Or: $$P_1T_2 = P_2T_1$$

Temperature must be converted to Kelvin!

Summary of 3 Gas Laws

Boyle's (Constant T): $P_1V_1 = P_2V_2$

Charles's (Constant P): $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ or $V_1T_2 = V_2T_1$

Gay-Lussac's (Constant V): $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ or $P_1T_2 = P_2T_1$

All of these laws refer to a specific quantity of gas.

Combined Gas Law

Three laws up to this point:

Boyle's Law (Constant Temperature):

$$P_1V_1 = P_2V_2$$

Charles' Law (Constant Pressure):

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

Gay-Lussac's Law (Constant Volume):

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

Combined Gas Law

A single expression that combines the three gas laws:

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

Or: $$P_1 V_1 T_2 = P_2 V_2 T_1$$

Temperature must be converted to Kelvin!

Combined Gas Law

Remembering that this equation is referring to a fixed amount of a gas, so

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} = k \text{ (a constant)}$$

Combined Gas Law Problem 1

A gas with a volume of 3.0 L at 40. °C has a pressure of 2.5 atm. What will the volume of this gas be at a pressure of 1.5 atm and a temperature of 90. °C?

P_1= 2.5 atm	V_1= 3.0L	T_1= 40. °C = 313 K	$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$
P_2= 1.5 atm	V_2= ?	T_2= 90. °C = 363 K	

$P_1V_1T_2 = P_2V_2T_1$ so $V_2 = \frac{P_1V_1T_2}{P_2T_1}$

$$V_2 = \frac{2.5\ \cancel{atm}\ x\ 3.0\ L\ x\ 363\ \cancel{K}}{1.5\ \cancel{atm}\ x\ 313\ \cancel{K}} = 5.7987 = 5.8\ L$$

Combined Gas Law Problem 2

A gas occupies 3.5 L at 600. mm pressure and 22 °C. What will the temperature of this gas be if the volume is changed to 1.8 L at 760 mm pressure?

P_1= 600. mm V_1= 3.5L T_1= 22 °C = 295 K

P_2= 760 mm V_2= 1.8 L T_2= ?

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$P_1V_1T_2 = P_2V_2T_1 \qquad T_2 = \frac{P_2V_2T_1}{P_1V_1}$$

$$T_2 = \frac{760\ \cancel{mm}\ x\ 1.8\ \cancel{L}\ x\ 295\ K}{600.\ \cancel{mm}\ x\ 3.5\ \cancel{L}} = \quad 192.17 = 190\ K$$

Ideal Gas Law

$$PV = nRT$$

Review: Combined Gas Law

Remembering that this equation is referring to a fixed amount of a gas, so

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} = k\,(\text{a constant})$$

$$k = nR$$

Avogadro's Law

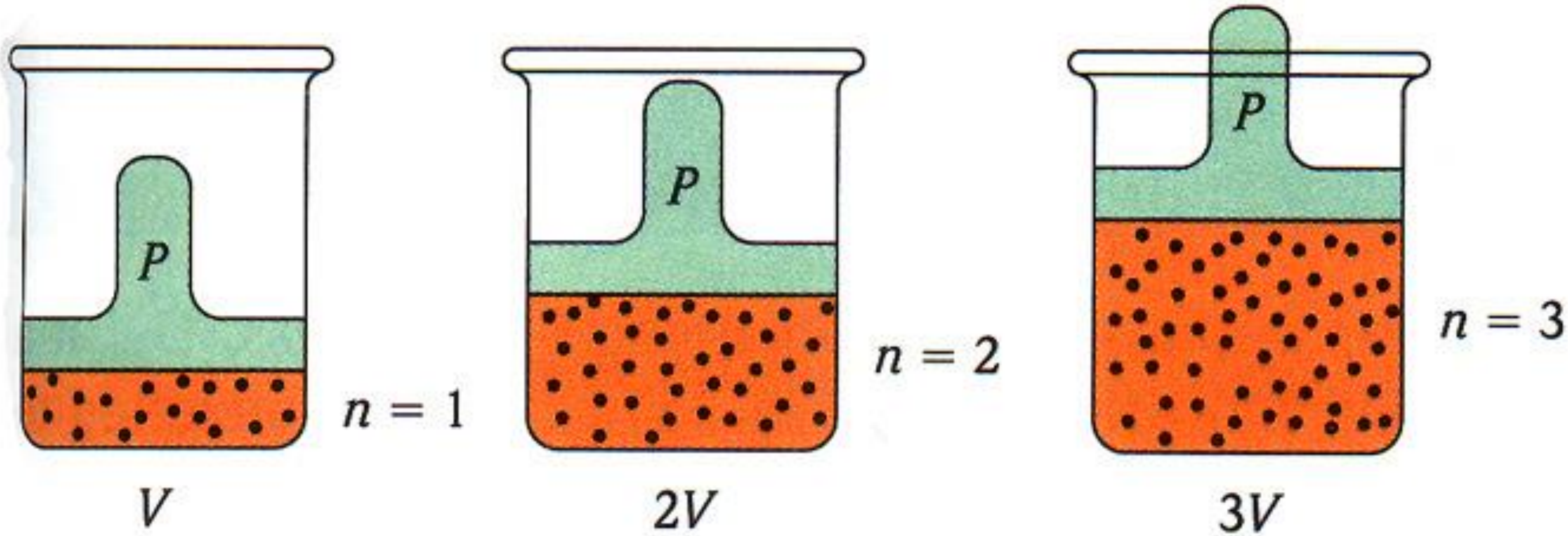
Avogadro investigated the relationship between the volume of a gas and the number of moles present in the gas sample.

Avogadro's Law

He found that as the number of moles is doubled, the volume doubles (*at constant temperature and pressure*).

$$\frac{V}{n} = a = constant$$

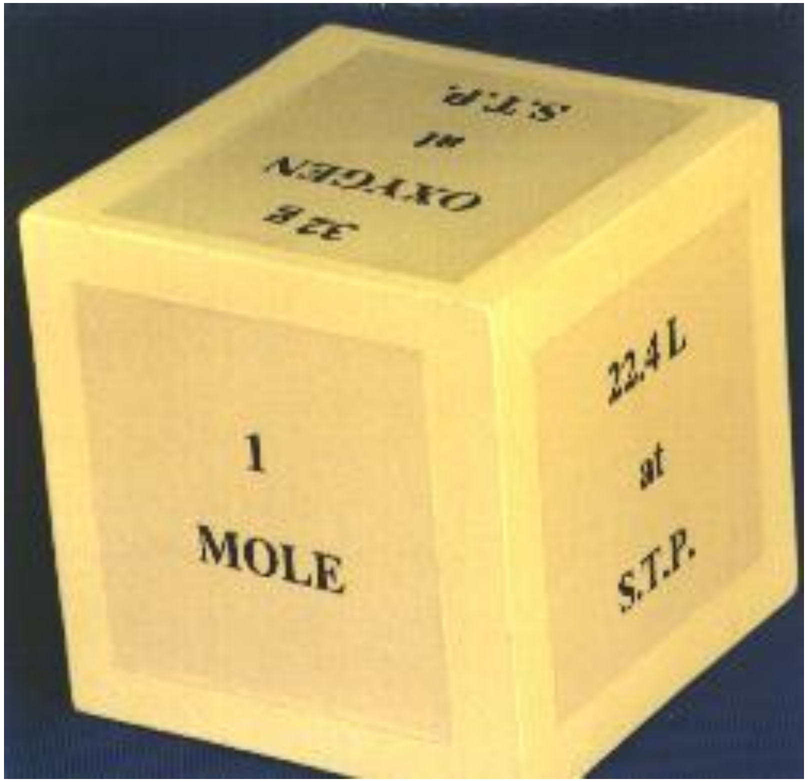
Avogadro's Law



The volume of a gas is directly proportional to the number of moles

Avogadro's Law:

Equal volumes of all gases at the same temperature and pressure contain the same number of molecules.
- Amedeo Avogadro



Ideal Gas Law

$$\frac{PV}{T} = constant = nR$$

$$or$$

$$PV = nRT$$

Ideal Gas Law

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

$$"constant" = nR$$

$$Where\ n = \#\ of\ moles\ \ and\ R = the\ gas\ constant$$

$$PV = nRT$$

Ideal Gas Law

$$PV = nRT$$

R is the ideal gas constant

$$R = 8.31\,\frac{L \cdot kPa}{mole \cdot K} \qquad R = .08206\,\frac{L \cdot atm}{mole \cdot K}$$

Ideal Gas Law Example

How many moles of gas are in a sample that has a pressure of 1.85 atm and a volume of 59.2 L at 37.0°C?

n = ?
P = 1.85 atm
V = 59.2 L
T = 310 K
R = .08206 L-atm/mol-K

$$PV = nRT \; so\!: \quad n = \frac{PV}{RT}$$

$$n = \frac{1.85\, a\cancel{tm}}{.08206 \frac{\cancel{L} \cdot a\cancel{tm}}{mol \cdot \cancel{K}}} \frac{59.2\, \cancel{L}}{310\cancel{K}} = 4.305\, mole$$

$$n = 4.21\, mole$$

Dalton's Law of Partial Pressure

$$P_{Total} = P_1 + P_2 + P_3 + \ldots\ldots\ldots$$

Dalton's Law of Partial Pressures

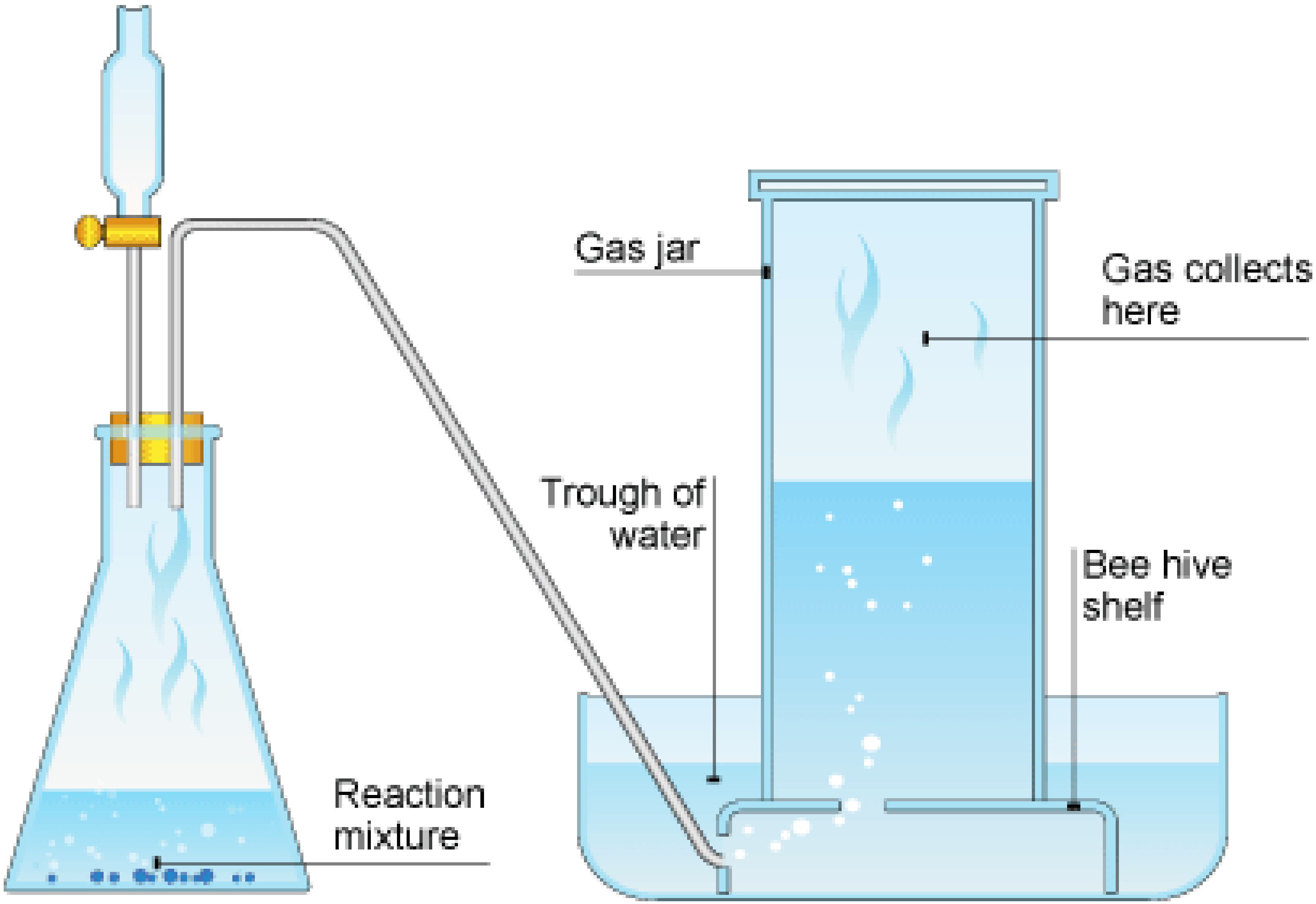
For a mixture of gases in a container

$$P_{\text{Total}} = P_1 + P_2 + P_3 + \ldots$$

This is particularly useful in calculating the pressure of gases collected over water.

Dalton’s Law of Partial Pressure

Gases are collected over water in the laboratory



Graham’s Law of Diffusion

Graham's Law of Diffusion

$$\frac{\text{rate of diffusion}_A}{\text{rate of diffusion}_B} = \frac{\sqrt{M_{\boldsymbol{B}}}}{\sqrt{M_{\boldsymbol{A}}}} = \frac{\sqrt{\text{density}_B}}{\sqrt{\text{density}_A}}$$