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Chapter 5
Gases

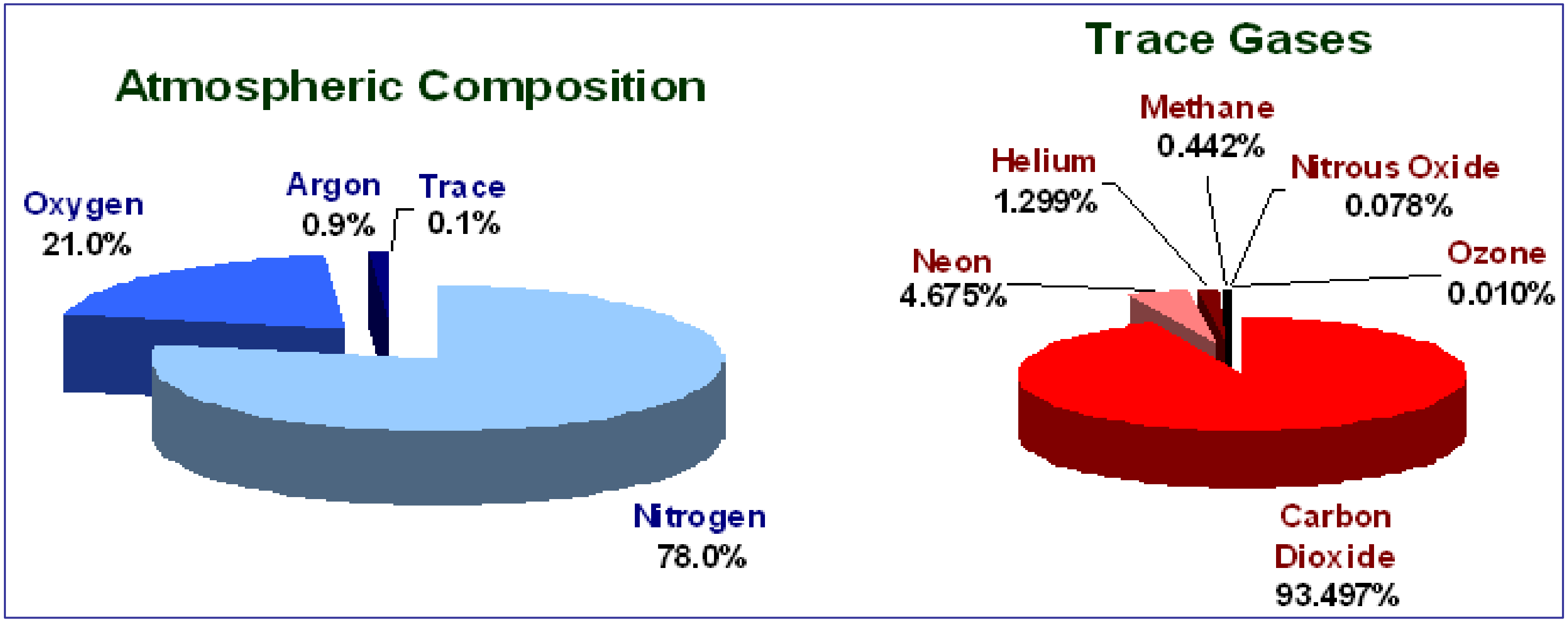
Ahmad Aqel Ifseisi
Assistant Professor of Analytical Chemistry
College of Science, Department of Chemistry
King Saud University
P.O. Box 2455 Riyadh 11451 Saudi Arabia
Building: 05, Office: 1A7 & AA53
Tel. 014674198, Fax: 014675992
Web site: http://fac.ksu.edu.sa/aifseisi
E-mail: ahmad3qel@yahoo.com
aifseisi@ksu.edu.sa

5.1
Substances that exist as gases

A **gas** is a substance that is normally in the gaseous state at ordinary temperatures and pressures.

A **vapor** is the gaseous form of any substance that is a liquid or a solid at normal temperatures and pressures.

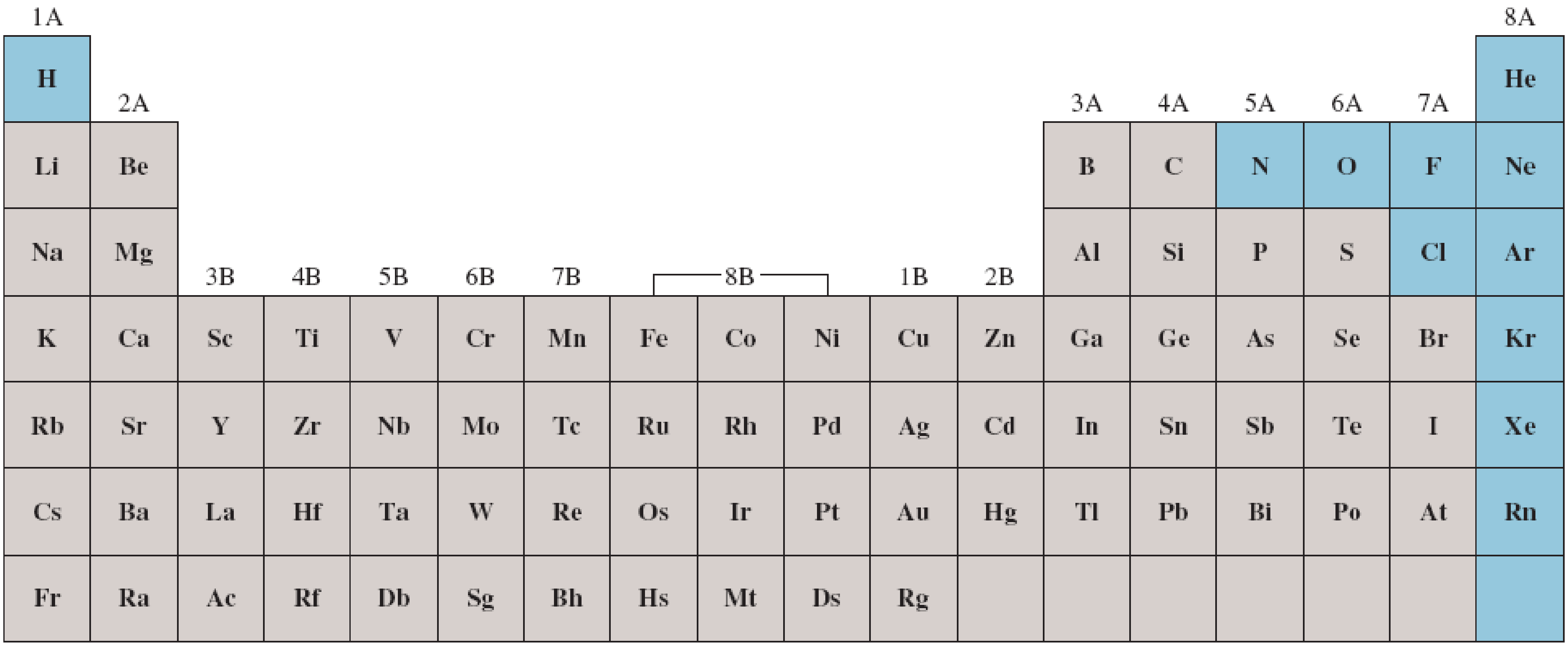
Thus, at 25^{o}C and 1 atm pressure, we speak of water vapor and oxygen gas.



Air is a complex mixture of several substances.

Elements that exist as gases at 25°C and 1 atm.

The noble gases (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	Compounds
H_2 (molecular hydrogen)	HF (hydrogen fluoride)
N_2 (molecular nitrogen)	HCl (hydrogen chloride)
O_2 (molecular oxygen)	HBr (hydrogen bromide)
O_3 (ozone)	HI (hydrogen iodide)
F_2 (molecular fluorine)	CO (carbon monoxide)
Cl_2 (molecular chlorine)	CO_2 (carbon dioxide)
He (helium)	NH_3 (ammonia)
Ne (neon)	NO (nitric oxide)
Ar (argon)	NO_2 (nitrogen dioxide)
Kr (krypton)	N_2O (nitrous oxide)
Xe (xenon)	SO_2 (sulfur dioxide)
Rn (radon)	H_2S (hydrogen sulfide)
	HCN (hydrogen cyanide)*

Characteristics of gases

All gases have the following physical characteristics:

- Gases assume the volume and shape of their containers.
- Gases are the most compressible of the states of matter.
- Gases can expand spontaneously to fill their containers.
- Gases will mix evenly and completely when confined to the same container.
- Gases have much lower densities than liquids and solids.
- Gases have relatively low molar masses.

Gases form homogeneous mixtures with each other regardless of the identities or relative proportions of the component gases.

5.2
Pressure of a gas

SI Units of Pressure

Pressure is one of the most readily measurable properties of a gas.

Pressure is a force applied per unit area

$$\text{pressure} = \frac{\text{force}}{\text{area}}$$

force = mass x acceleration

the SI unit of force is the **newton (N),** where
1 N = 1 kg m/s^2

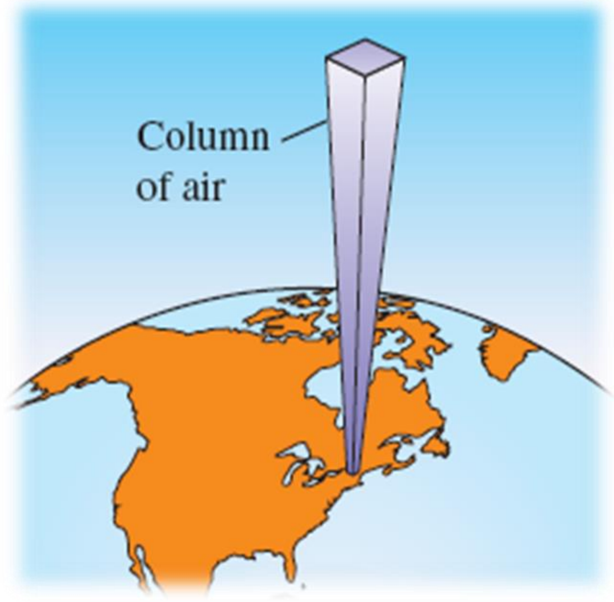
The SI unit of pressure is the **pascal (Pa),** defined as one N per m^2:
1 Pa = 1 N/m^2

Atmospheric Pressure

The force experienced by any area exposed to Earth's atmosphere is equal to the weight of the column of air above it.

Atmospheric pressure is the pressure exerted by Earth's atmosphere .

The actual value of atmospheric pressure depends on location, temperature, and weather conditions.



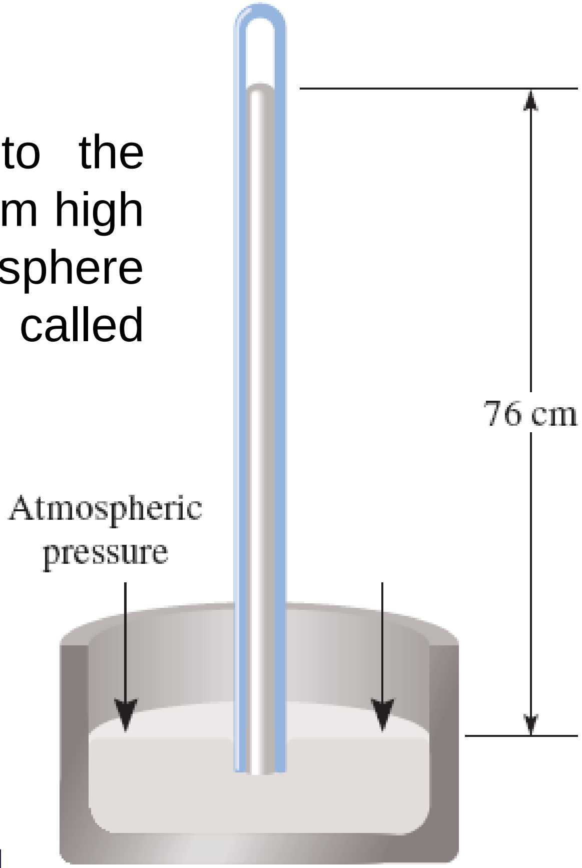
A column of air extending from sea level to the upper atmosphere.

How is atmospheric pressure measured?

The **barometer** is the most familiar instrument for measuring atmospheric pressure; consists of a long glass tube, closed at one end and filled with mercury. If the tube is inverted in a dish of mercury so that no air enters the tube, some mercury will flow out of the tube into the dish, creating a vacuum at the top. The weight of the mercury remaining in the tube is supported by atmospheric pressure acting on the surface of the mercury in the dish.

Standard atmospheric pressure (1 atm) is equal to the pressure that supports a column of mercury exactly 760 mm high at 0°C at sea level. In other words, the standard atmosphere equals a pressure of 760 mmHg. The mmHg unit is also called the torr.

1 torr = 1 mmHg
1 atm = 760 mmHg



Units of Pressure

Pressure Units						
	pascal (Pa)	**bar (bar)**	**atmosphere (atm)**	**torr (torr)**	**pound-force per square inch (psi)**	**kilogram-force per square centimeter (kgf/cm^2)**
1 Pa	≡ 1 N/m^2	10^{-5}	9.8692×10^{-6}	7.5006×10^{-3}	145.04×10^{-6}	1.01972×10^{-5}
1 bar	100,000	≡ 10^6 dyn/cm^2	0.98692	750.06	14.504	1.01972
1 atm	101,325	1.01325	≡ 1 atm	760	14.696	1.03323
1 torr	133.322	1.3332×10^{-3}	1.3158×10^{-3}	≡ 1 torr ≈ 1 mmHg	19.337×10^{-3}	1.35951×10^{-3}
1 psi	6,894.76	68.948×10^{-3}	68.046×10^{-3}	51.715	≡ 1 lbf/in^2	7.03059×10^{-2}
1 kgf/cm^2	98,066.5	0.980665	0.967838	735.5576	14.22357	≡ 1 kgf/cm^2

Example reading: 1 Pa = 1 N/m^2 = 10^{-5} bar = 9.8692×10^{-6} atm = 7.5006×10^{-3} torr, etc.

Note: mmHg is an abbreviation for millimetre of mercury

$$1\text{ atm} = 760\text{ mm Hg} = 760\text{ torr} = 1.01325\times10^5\text{ Pa} = 101.325\text{ kPa}$$

EXAMPLE

The pressure outside a jet plane flying at high altitude falls considerably below standard atmospheric pressure. Therefore, the air inside the cabin must be pressurized to protect the passengers. What is the pressure in atmospheres in the cabin if the barometer reading is 688 mmHg?

1 atm = 760 mmHg,

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

$$\text{pressure} = 688 \cancel{\text{mmHg}} \times \frac{1 \text{ atm}}{760 \cancel{\text{mmHg}}} = 0.905 \text{ atm}$$

Practice Exercise

Convert 749 mmHg to atmospheres.

EXAMPLE

The atmospheric pressure in San Francisco on a certain day was 732 mmHg. What was the pressure in kPa?

1 atm = 1.01325 x 10^5 Pa = 760 mmHg

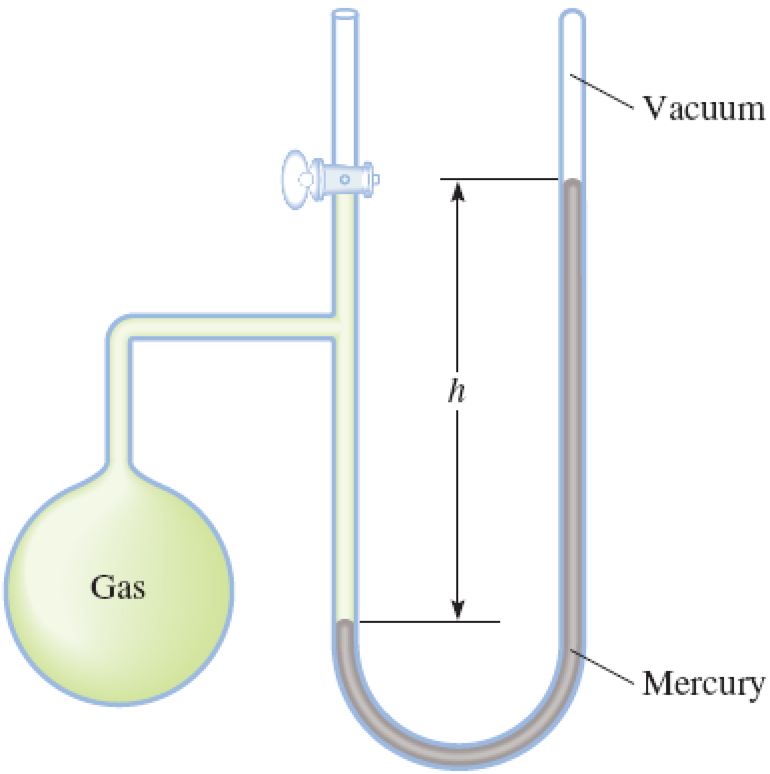
$$\frac{1.01325 \times 10^5 \text{ Pa}}{760 \text{ mmHg}}$$

$$\begin{aligned} \text{pressure} &= 732 \text{ mmHg} \times \frac{1.01325 \times 10^5 \text{ Pa}}{760 \text{ mmHg}} \\ &= 9.76 \times 10^4 \text{ Pa} \\ &= 97.6 \text{ kPa} \end{aligned}$$

Practice Exercise

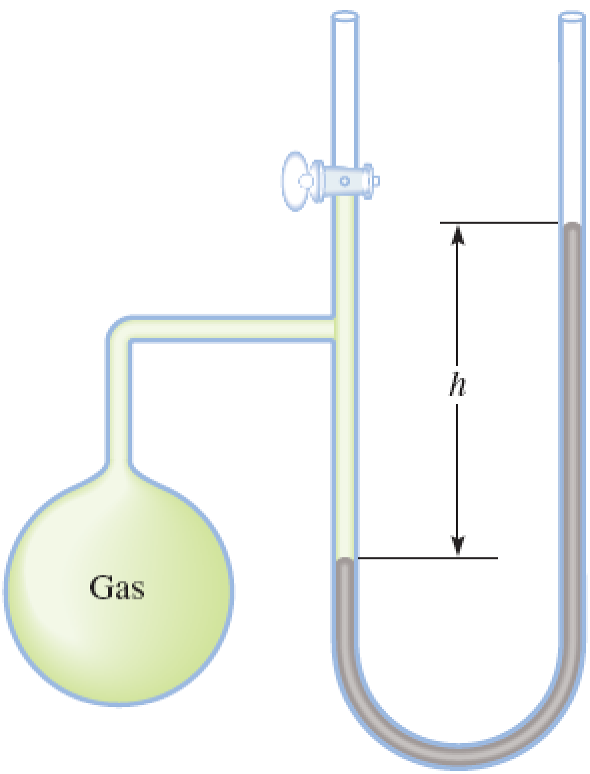
Convert 295 mmHg to kilopascals.

A ***manometer*** is *a device used to measure the pressure of gases other than the atmosphere.*



$P_{gas} = P_h$

The *closed-tube manometer* is normally used to measure pressures below atmospheric pressure.



$P_{gas} = P_h + P_{atm}$

The *open-tube manometer* is better suited for measuring pressures equal to or greater than atmospheric pressure.

5.3
The gas laws

Experiments with a large number of gases reveal that four variables are needed to define the physical condition, or state of a gas:

- Temperature (**T**)
- Pressure (**P**)
- Volume (**V**)
- Amount of gas, usually expressed as the number of moles (**n**).

The equations that express the relationships among **T**, **P**, **V**, and **n** are known as the **gas laws**.

-If we want to study the relationship between **T** & **P** we should fix **V** & **n** constant.
-If we want to study the relationship between **P** & **V** we should fix **T** & **n** constant.
and so on,,,

The Pressure-Volume Relationship: Boyle's Law

The pressure of a fixed amount of gas at a constant temperature is inversely proportional to the volume of the gas.

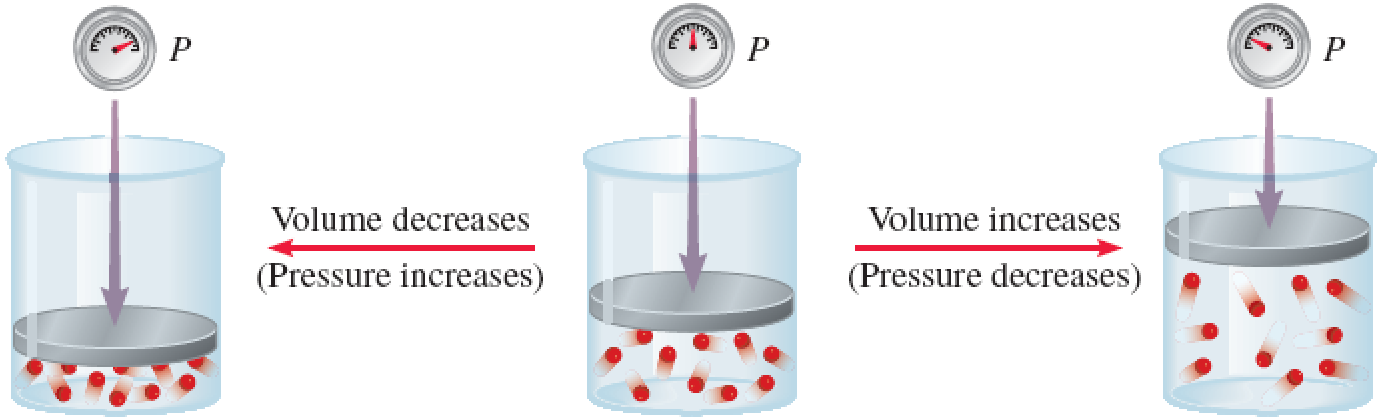
A mathematical expression (inverse relationship between **P** and **V**):

$$P \propto \frac{1}{V} \qquad P = k_1 \times \frac{1}{V} \qquad PV = k_1$$

$$P_1V_1 = k_1 = P_2V_2$$

$$\boxed{P_1V_1 = P_2V_2}$$

where k_1 is the proportionality constant, V_1 & V_2 are the volumes at pressures P_1 & P_2.
The product of **P** & **V** of a gas at constant **T** and amount of gas is a constant.
P times **V** is always equal to the same constant; the value of the constant depends on **T** & **n**.

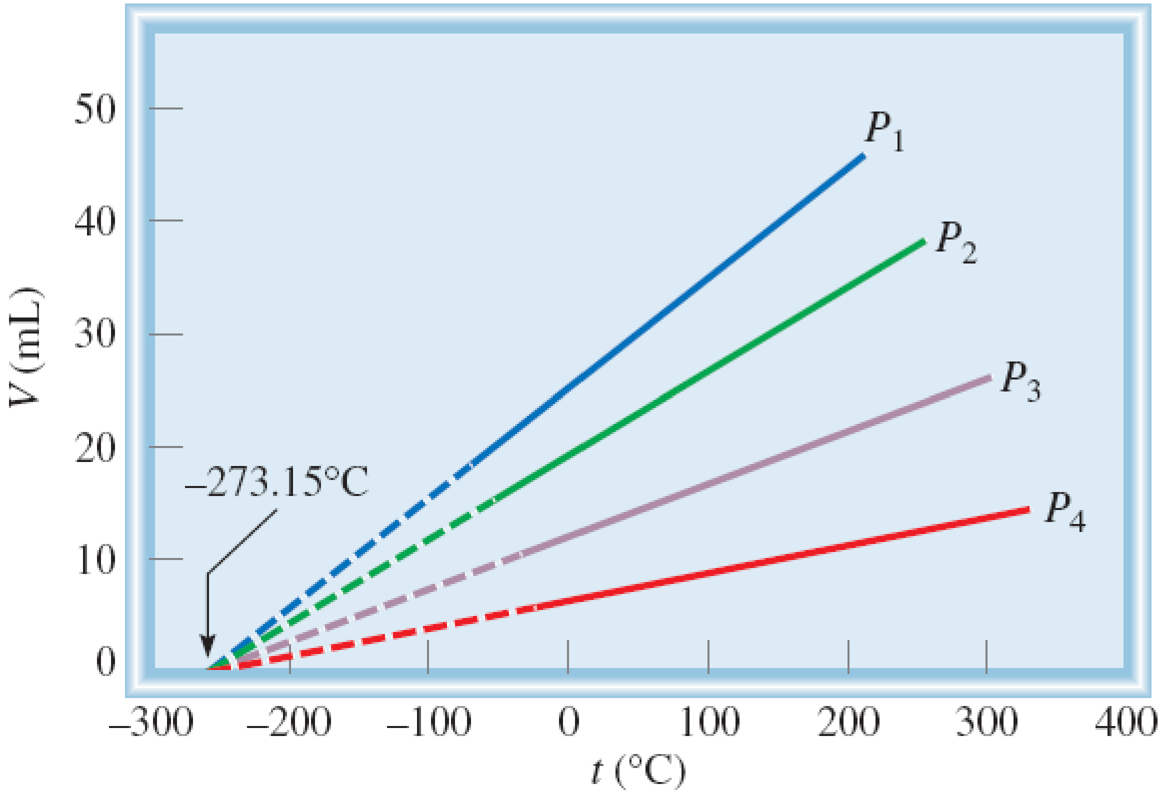


Increasing or decreasing the volume of a gas at a constant temperature.

The Temperature-Volume Relationship: Charles's and Gay-Lussac's Law

At constant pressure, the volume of a gas sample expands when heated and contracts when cooled.

At any given pressure, the plot of volume versus temperature yields a straight line. By extending the line to zero volume, we find the intercept on the temperature axis to be -273.15ºC.



In 1848 Lord Kelvin identified -273.15ºC as **absolute zero**, theoretically the lowest attainable temperature. Then he set up an **absolute temperature scale**, now called the **Kelvin temperature scale**.

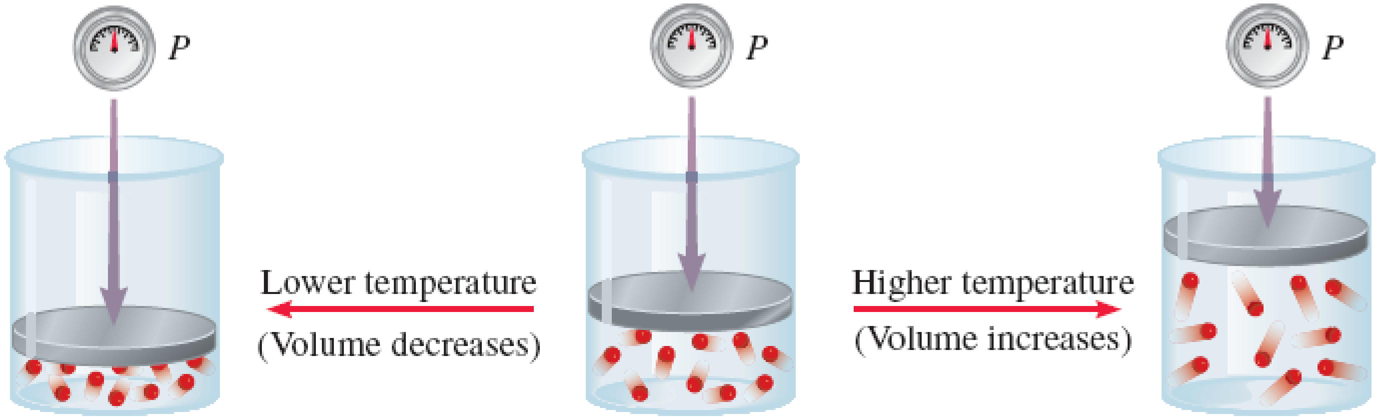
The volume of a fixed amount of gas maintained at constant pressure is directly proportional to the absolute temperature of the gas.

$$V \propto T \qquad V = k_2 T \qquad \frac{V}{T} = k_2$$

where k_2 is the proportionality constant. The equation is known as **Charles's and Gay-Lussac's law**, or simply **Charles's law**.

$$\frac{V_1}{T_1} = k_2 = \frac{V_2}{T_2} \qquad \boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$$

where V_1 and V_2 are the volumes of the gas at temperatures T_1 and T_2 (both in kelvins), respectively.



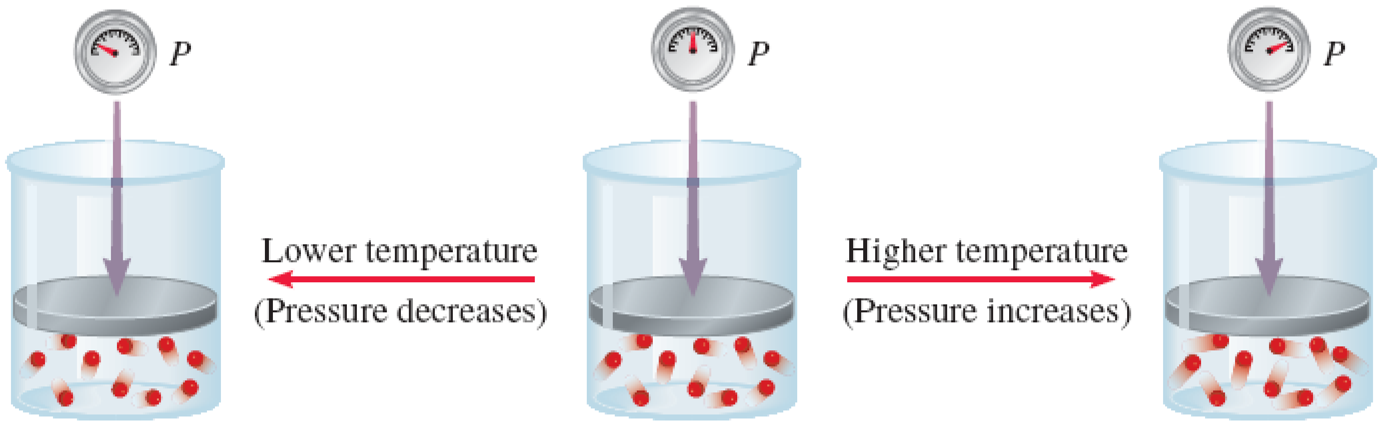
Heating or cooling a gas at constant pressure

Another form of Charles's law shows that at constant amount of gas and volume, the pressure of a gas is proportional to temperature

$$P \propto T \qquad P = k_3 T \qquad \frac{P}{T} = k_3$$

$$\frac{P_1}{T_1} = k_3 = \frac{P_2}{T_2} \qquad \boxed{\frac{P_1}{T_1} = \frac{P_2}{T_2}}$$

where P_1 and P_2 are the pressures of the gas at temperatures T_1 and T_2, respectively.



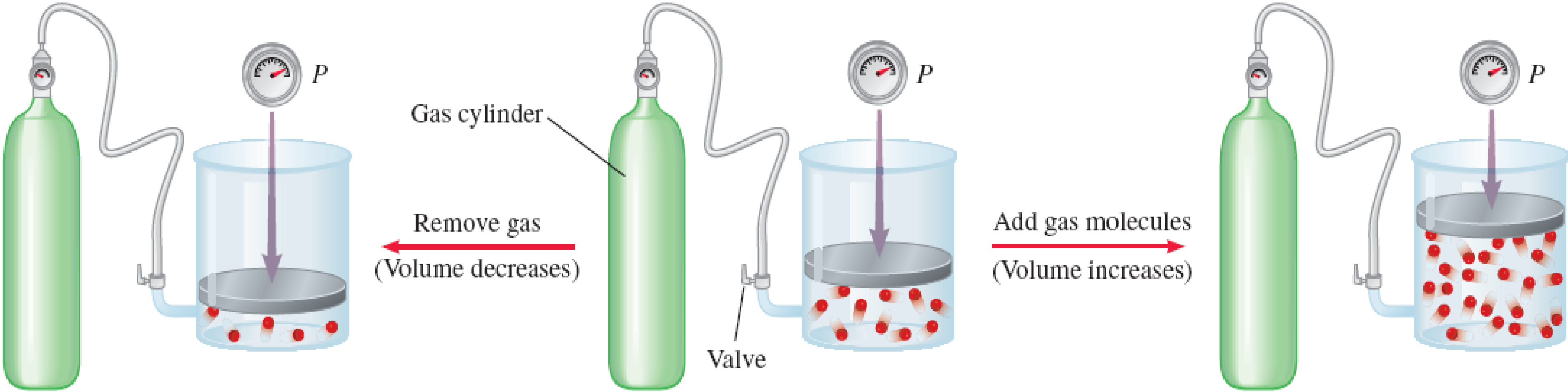
Heating or cooling a gas at constant volume

The Volume-Amount Relationship: Avogadro's Law

Avogadro stating that at the same **T** and **P**, equal volumes of different gases contain the same number of molecules (or atoms if the gas is monatomic).
The volume of any given gas must be proportional to the number of moles of molecules present:

$$V \propto n \qquad \boxed{V = k_4 n}$$

where **n** represents the number of moles, k_4 is the proportionality constant. The equation is the mathematical expression of **Avogadro's law**; states that at constant **P** & **T**, the volume of a gas is directly proportional to the number of moles of the gas present.



Dependence of volume on amount of gas at constant temperature and pressure

According to Avogadro's law; when two gases react with each other, their reacting volumes have a simple ratio to each other. If the product is a gas, its volume is related to the volume of the reactants by a simple ratio. e.g., consider the synthesis of ammonia from molecular hydrogen and molecular nitrogen:

$$3H_2(g) + N_2(g) \longrightarrow 2NH_3(g)$$

3 moles	1 mole	2 moles
3 molecules	1 molecule	2 molecules

Because, at the same **T** & **P**, the volumes of gases are directly proportional to the number of moles of the gases present, we can now write:

$$3H_2(g) + N_2(g) \longrightarrow 2NH_3(g)$$

3 volumes	1 volume	2 volumes

The volume ratio of molecular hydrogen to molecular nitrogen is 3:1, and that of ammonia (the product) to the sum of the volumes of molecular hydrogen and molecular nitrogen (the reactants) is 2:4 or 1:2


Thank You!