CHEM 1300
Week 5 - 8

Lectures 5 to 8 (Week 5 to 8):
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Chapter 10: Gases
Chapter 14: Chemical Kinetics
Chapter 15: Chemical Equilibrium
Chapter 5: Thermochemistry

Gases
(Chapter 10)

1. **Pressure**
2. **Gas Laws**
3. **Ideal gas equation**
4. **Kinetic-molecular theory of gases**
5. **Molecular diffusion**

Characteristics of Gases

- **The atmosphere is composed of gaseous mixture of oxygen and nitrogen – air.**

- **Unlike liquids and solids, gases**
  - **expand to fill in the containers.**
  - **are highly compressible.**
  - **have relatively low densities.**

- **Two (or more) gases mix together to form a homogenous mixture.**

Pressure

- **Gases exert a pressure on any contacting surface.**

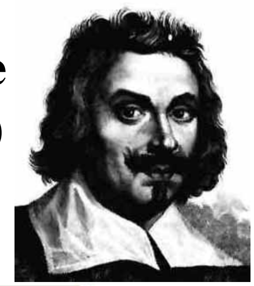
- **Pressure is defined as the amount of force applied to the area of surface.**

$$[\mathrm{N\ m^{-2}}] \quad \boxed{P = \frac{F}{A}} \quad \begin{matrix}[\mathrm{N}] \\ [\mathrm{m^2}]\end{matrix}$$

- **Units of pressure**
  - **Pascal : 1 Pa = 1 N m^{-2}**
  - **Bar: 1 bar = 10^5 Pa = 100 kPa**

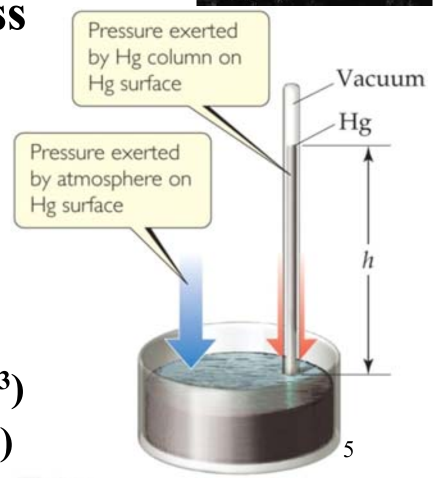
Measurements of pressure

- **Barometer was invented to measure the atmospheric pressure by the heights (*h*) of mercury columns by Torricelli.**



- **In atmosphere, an inverted glass tube, completely filled with mercury (Hg), exerts a column height of 76 cm on a dish of mercury.**

$$P = h\rho g$$



ρ – density of mercury (13.59×10^3 kg m^{-3})
g – acceleration due to gravity (9.81 m s^{-2})

- **The mass of mercury column is balanced by the external pressure:**

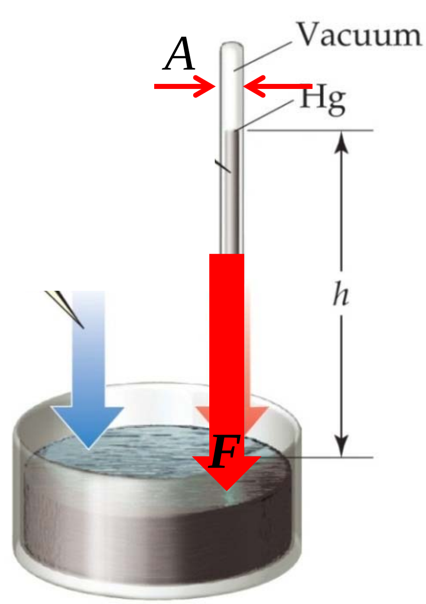
$$F = mg$$

$$F = V\rho g$$

$$F = hA\rho g$$

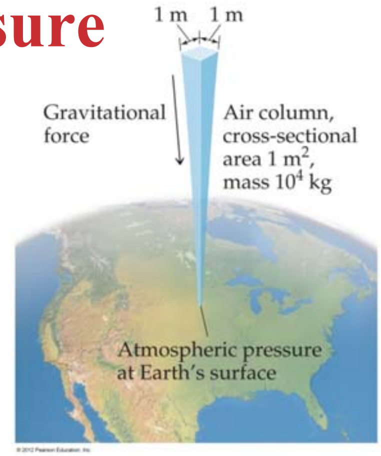
$$P = \frac{F}{A} = \frac{h\cancel{A}\rho g}{\cancel{A}}$$

$$P = h\rho g$$



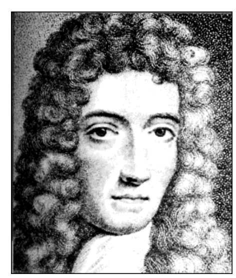
m **– mass of mercury column**
A **– cross sectional area of column**

Atmospheric pressure



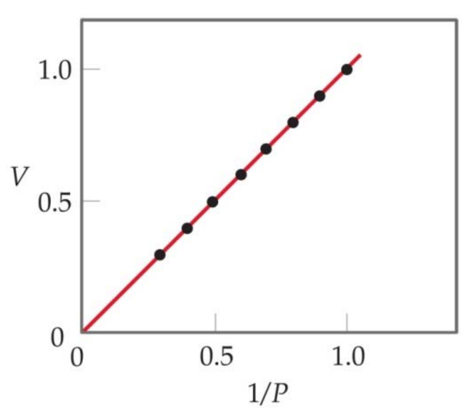
- **One atmospheric pressure (atm) is equal to:**
  - **760 mm Hg**
  - **760 Torr**
  - **101325 Pa or N m^{-2}**
  - **1.01325 bar**
- **An air column (cross-sectional area ≈ 1 m^2) of mass 10,000 kg, extends through the entire atomsphere on earth.**
- **Blood pressure is expressed in the unit of mmHg or Torr.**

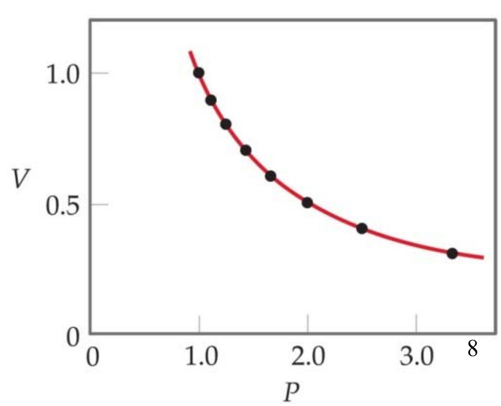
The pressure – volume relationship: Boyle's Law



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

$$V = \text{constant} \times \frac{1}{P} \quad \text{or} \quad PV = \text{constant}$$



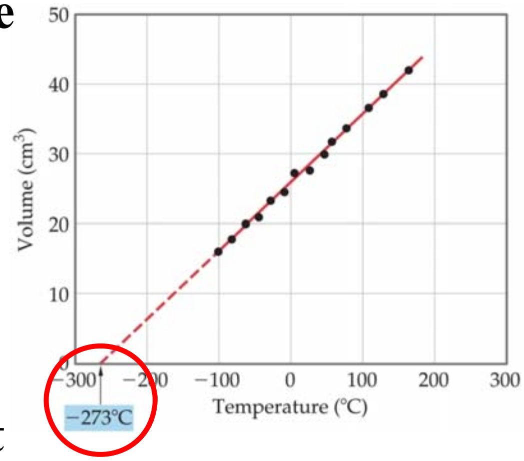


The temperature – volume relationship: Charles's Law

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

$$V \propto T$$

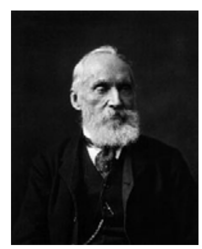
- A plot of V versus T is a straight line.



- Extrapolation of $V - T$ plot gives to the T of –273 °C.

$$\boxed{T\,(K) = T\,(^oC) + 273.15}$$

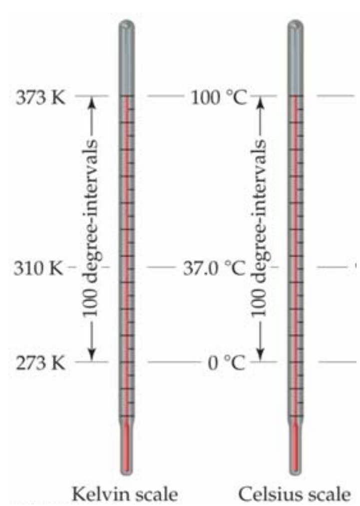
Temperature scale



- The absolute temperature scale was proposed by William Thomson, 1st Baron Kelvin in 1848.

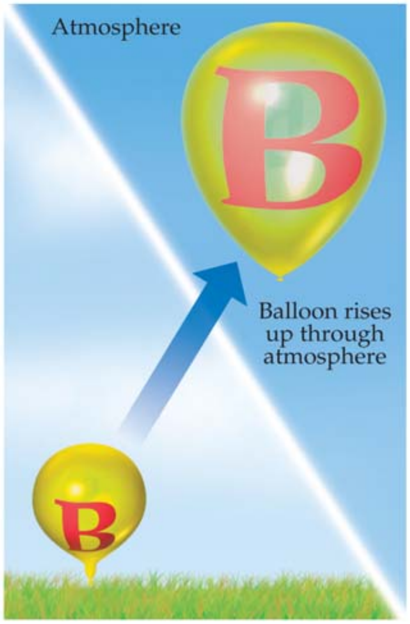
$$T\,(K) = T\,(^oC) + 273.15$$

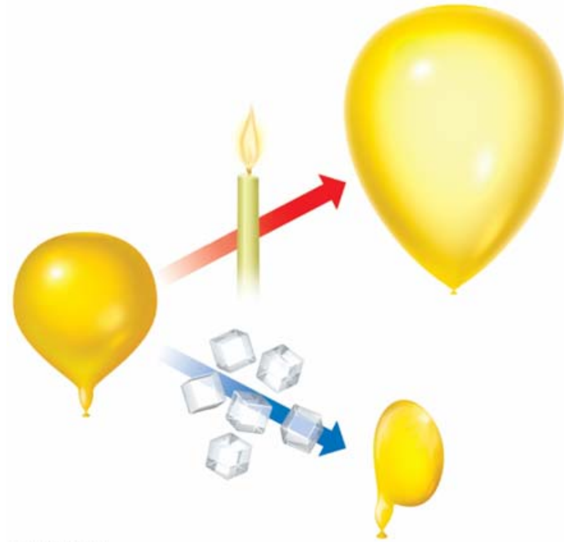
- On this scale, 0 K called ***absolute zero***, equals –273.15 °C.



$T\,(K)$ $T\,(^oC)$

Boyle's Law and Charles's Law





Does atmospheric pressure increase or decrease as altitude goes high?

Balloon changes its volume when it is heated up or cooled down.

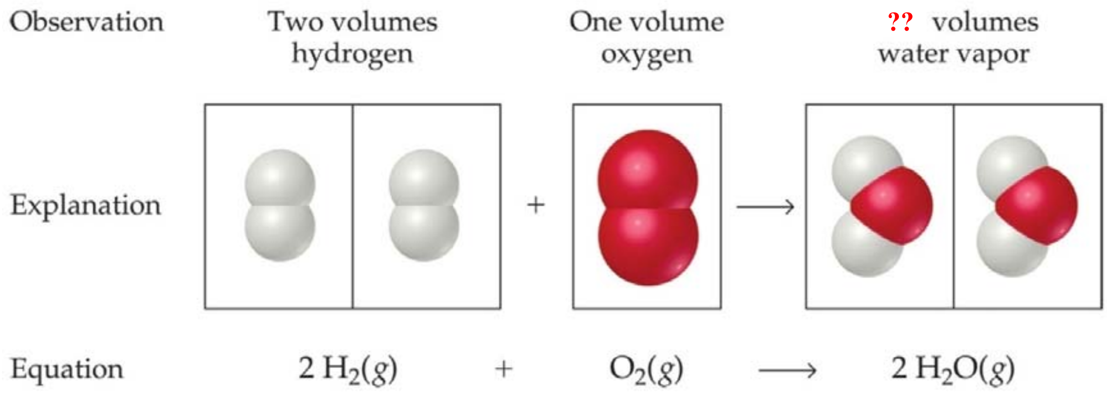
The quantity – volume relationship: Avogadro's Law

- **The volume of a gas at constant temperature and pressure is directly proportional to the number of moles (*n*) of the gas.**
- **Mathematically, this means** $V = \text{constant} \times n$

<u>Avogardro's hypothesis</u>

Equal volume of gases at the same *T* and *P* contain equal numbers of molecules.

	He	N_2	CH_4
Volume	22.4 L	22.4 L	22.4 L
Pressure	1 atm	1 atm	1 atm
Temperature	0 °C	0 °C	0 °C
Mass of gas	4.00 g	28.0 g	16.0 g
Number of gas molecules	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}



Describe what happens to the volume of gases after the reaction.

Describe what happens to gas cylinder (with a movable piston) subjected to following changes:

Changes	Comments
(A) Heating the gas at constant pressure	
(B) Reducing the volume at constant temperature	
(C) Injecting additional gas, keeping temperature and volume constant	

Ideal-Gas Equation

Boyle's law: $V \propto \frac{1}{P}$

Charles's law: $V \propto T$

Avogadro's law: $V \propto n$

- **Combining all three equations, we get**

$$V \propto \frac{nT}{P} \quad \text{or} \quad V = \underbrace{\text{constant}}_{R} \times \frac{nT}{P}$$

Universal gas constant

- **The constant of proportionality is known as *R*, the gas constant.**

$$V = R\frac{nT}{P}$$

$$\boxed{PV = nRT}$$

Units	Numerical Value
L-atm/mol-K	0.08206
J/mol-K*	8.314
cal/mol-K	1.987
m^3-Pa/mol-K*	8.314
L-torr/mol-K	62.36

*SI unit.

- **The value of *R* is determined by the measurement that 1 mole of an ideal gas occupies a volume of 22.414 liter at 1 atm and 0 °C (273.15 K).**
- **The condition of 1 atm and 0 °C is referred as standard temperature and pressure (STP).**

Ideal gas law in term of gas density

$$PV = nRT$$

mass

$$PV = \frac{m}{M} RT$$

$$\frac{m}{V} = \frac{PM}{RT}$$

molar mass

CO_2 (as fire extinguisher agent) is denser than air.

Mass density of the gas:

$$d = \frac{m}{V} = \frac{PM}{RT}$$

Hot-air balloon

Helium balloon

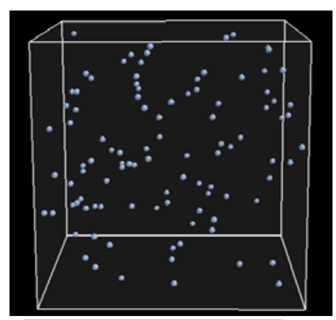
Kinetic-Molecular theory of gases

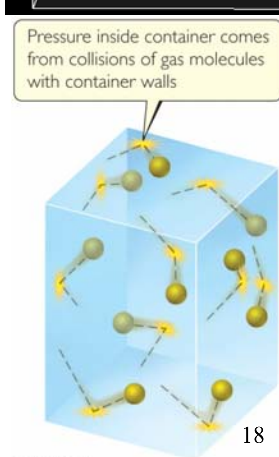
It states that:

1. Gases consist of large numbers of molecules that are in continuous and random motion.

2. The occupied volume of all gaseous molecules is negligible.

3. Attractive and repulsive forces between gas molecules are negligible.





Kinetic-Molecular theory of gases

4. Energy is transferred between molecules during collisions.

The *average* kinetic energy (K.E.) of the molecules is proportional to the absolute temperature of gas.

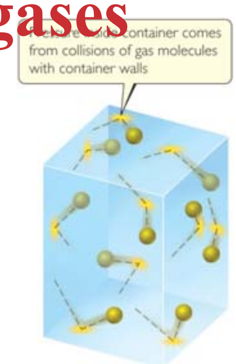
$$\text{Average} \quad \text{K.E.} = \frac{3}{2} kT$$

where *k* is Boltzmann constant.

(1.381×10^{-23} J K^{-1})

At a given temperature, all gas molecules in a mixture have the same average kinetic energy, independent of molecular masses.

Kinetic-Molecular theory of gases



The theory explains both pressure and temperature at molecular level:

- **The pressure of a gas is caused by collisions of atoms/molecules with the container walls.**
  - **A smaller box experiences more collisions by gases.**
  - $V\downarrow \Rightarrow P\uparrow$
- **The absolute temperature of a gas is a measure of average kinetic energy.**
  - **A higher temperature allows molecules to move faster and make more collisions.**
  - **$T\uparrow$ at constant $V \Rightarrow P\uparrow$.**

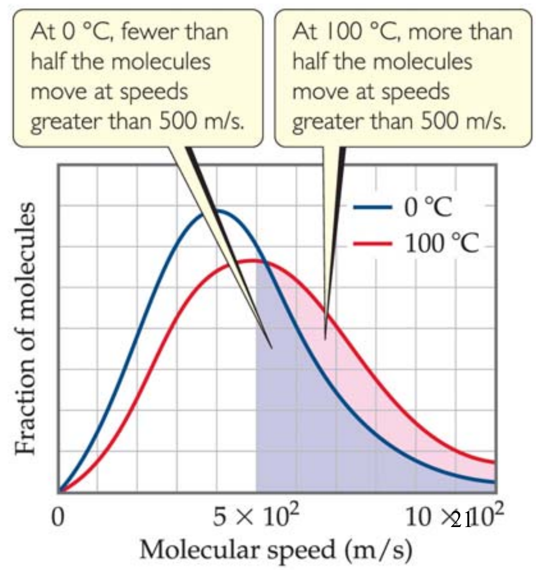
Distribution of molecular speed

- **Molecules make collisions, the individual molecules move at different speeds – speed distribution.**

$$\text{Average}\quad \text{K.E.} = \frac{1}{2} m u^2$$

$$\Rightarrow u = \sqrt{\frac{3kT}{m}} = \sqrt{\frac{3RT}{M}}$$

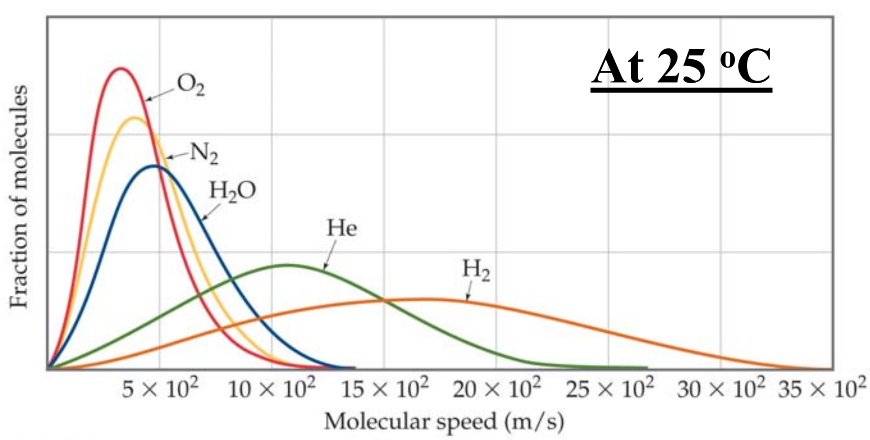
m **is mass of molecule.**
u **is the root-mean-square (*rms*) speed.**



RMS speed

- **Different gases have different *rms* speeds at a given temperature.**
- **The lighter the molecule, the faster it moves, the higher its *rms* speed.**

$$u = \sqrt{\frac{3RT}{M}}$$



Molecular diffusion

- **Molecules are in constant and random motion and tend to move from regions where they are in higher concentration to regions where they are less concentrated – Diffusion.**

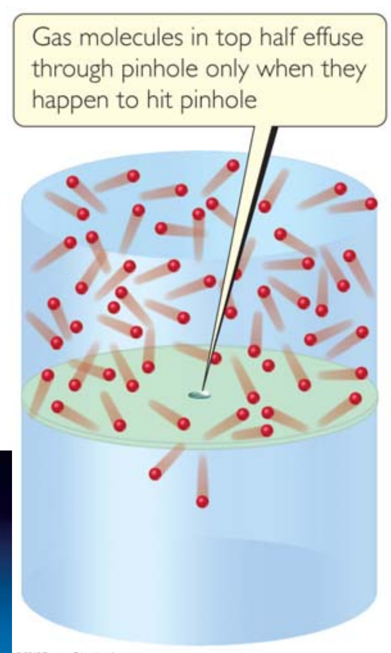
- **Smell of perfume – the perfume molecules in the bottle diffuse and mix with the molecules of air in the room.**

E.g. Diffusion of Bromine
http://www.youtube.com/watch?v=_oLPBnhOCjM

Molecular effusion

- **The escape of molecule through a tiny hole – effusion.**

- **Smell of food comes out from a sealed plastic bag or wrapper through some invisible tiny holes.**



Graham's law

- **Diffusion / effusion rate (*r*) depends on the speed of molecules :**

$$r \propto u = \sqrt{\frac{3kT}{m}}$$

- **At same *T*, the ratio of diffusion rate between two gases of different masses is:**

$$\frac{r_1}{r_2} = \sqrt{\frac{m_2}{m_1}}$$

where m_1 and m_2 is the masses of gases 1 and 2.

- **Gas of lighter mass diffuse faster than gas of heavier mass.**