INTRODUCTION TO CHEMISTRY
(SC-230)

UNIT 8. GASES

1. Gas pressure and its measurement.
2. Boyles Law
3. Charles Law
4. Avogadro's Law
5. The Ideal Gas Law.
6. Law of partial pressure

UNIT 8. GASES

- Gases are composed of small molecules, and they have several characteristics that distinguish them from liquids and solids.
  - You can compress gases into smaller and smaller volumes
  - Gases, unlike for liquids and solids, can relate **pressure** (P), **volume** (V), **temperature** (T), and **molar amount** (*n*) of substance by the ideal gas law.

1. Gas Pressure and its Measurement

- **Pressure** = *the force exerted per unit area of surface*

Example: A person standing on a table
  - The person is exerting a force → This force is a downward pressure on the table → The force is a result of gravity → Air is also exerting additional pressure on the table

- We can calculate the force on the table by using the following equation:

  Force = mass X constant acceleration of gravity (9.81 m/s^2)

- We can then calculate the pressure on the table by using the following equation:

  Pressure = Force / Area

Example: Calculate the pressure of a new penny on a table

Mass = 2.5×10^{-3} kg
Constant acceleration of gravity = 9.81 m/s^2
Area of a penny = 2.7×10^{-4} m^2

1) **Force = mass X constant acceleration of gravity**
   = 2.5×10^{-3} kg x 9.81 m/s^2 = **2.5×10^{-2} kg · m/s^2**

2) **Pressure = Force / Area**
   Pressure = (2.5×10^{-2} kg m/s^2) / 2.7×10^{-4} m^2 = 92.6 kg/(m·s^2)

Measuring gas pressure
  - The unit for pressure is **kg/(m · s^2)** . This is given the name **Pascal (Pa).**

Example: The pressure exerted by the coin was approximate 100 Pa. whereas the pressure exerted by the atmosphere is approximately 1000 times larger (100,000 Pa). The Pascal is an extremely small unit.

- **Barometer =** *a device for measuring the pressure of the atmosphere***.**
    - Units used to measure atmospheric pressure: **millimeters of mercury (mmHg)** (called the **torr**)
    - **1 atmosphere (atm) = 760 mmHg (at sea level) or 760 torr**
- **Manometer** = *a device for measuring the pressure of a gas or liquid in a vessel.*

Table 5.2 Important Units of Pressure	
Unit	**Relationship or Definition**
Pascal (Pa)	$kg/(m \cdot s^2)$
Atmosphere (atm)	$1\ atm = 1.01325 \times 10^5\ Pa \simeq 101\ kPa$
mmHg, or torr	$760\ mmHg = 1\ atm$
Bar	$1.01325\ bar = 1\ atm$

Gas laws

- All gases behave simply in terms of temperature, pressure, volume and molar amount. Therefore a number of gas laws have been made to clarify the behavior of gases.
- We are going to look at the following gas laws.

1) Boyle's law
2) Charles Law
3) Avogadro's Law
4) The Ideal Gas Law.
5) Dalton's Law of partial pressure

2. Boyle's Law (Pressure and Volume) [Note: Temperature is constant]

- Compared to liquids and solids which are relatively uncompressible, gases are **compressible** (we can squish it)
- The compressibility of gases was first studied by Robert Boyle in 1661.
  - Boyle noticed that:
    - When he poured mercury into the open end of a J-shaped tube the volume of enclosed gas decreased.

- From this he came up with Boyle's Law

Boyles Law = *the volume of a sample of gas at a given temperature varies inversely with the applied pressure e.g., if the pressure is doubled the volume is halved***.**

$V = 1/P$ [V=volume; P=pressure – they are inversely proportional]

3. Charles Law (Temperature and Volume) [Note: Pressure is constant]

- Temperature also effects gas volume.
  - A gas contracts when it cools and expands when heated.

- Effect of temperature on gas was first observed by Jacques Alexandre Charles in 1787. Charles was a pioneer of hot air balloons and hydrogen filled balloons.

Charles law = *the volume occupied by any sample of gas at a constant pressure is directly proportional to the <u>absolute</u> temperature e.g., doubling the temperature of a gas doubles its volume.*

$V = bT$ [V=volume; T=temperature (in Kelvin); b=constant]

COMBINE GAS LAW – BOYLES AND CHARLES COMBINED

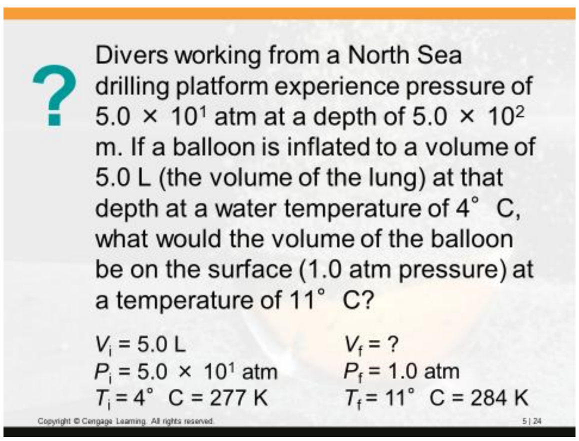
- The volume of a sample of gas at constant pressure is inversely proportional to the pressure and directly proportional to the absolute temperature.
- The mathematical relationship:

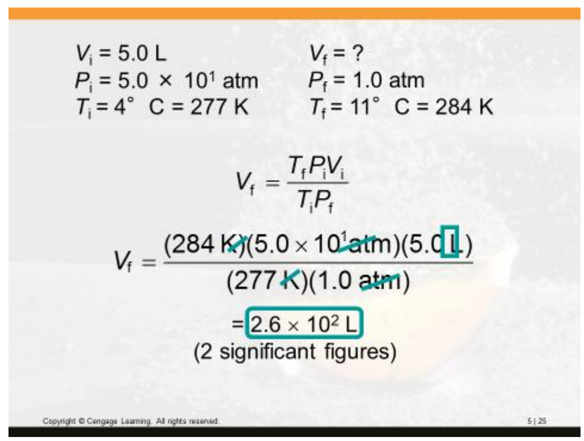
$$V \propto \frac{T}{P}$$

- In equation form: $\mathbf{(P_1V_1)/T_1 = (P_2V_2)/T_2}$ **or**

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

$$\frac{P_iV_i}{T_i} = \frac{P_fV_f}{T_f}$$





4. Avogadro's Law (Molar amount and Volume)

- Volume and amount can also be described by a law.
- This was initiated by Lussac (1778-1850). Lussac noticed that in experiments of gas reactions the volume of the reactant gases were in ratios of small whole numbers.
  - **<u>Example:</u>** 2 volumes of hydrogen react with 1 volume of oxygen.

- 3 years later, Avogadro created his law.

> **Avogadro's law =** *equal volumes of any 2 gases at the same temperature and pressure contain the same number of molecules e.g., 2 volumes of hydrogen contains twice as many molecules as 1 volume of oxygen. This is the case for any gas.*

- Therefore 1 mole of every gas contains the same number of molecules.
- This number is Avogadro's number = $\mathbf{6.022 \times 10^{23}}$
  - Therefore 1 mole of Oxygen contains 6.022×10^{23} molecules.
  - Therefore 1 mole of Hydrogen contains 6.022×10^{23} molecules.
  - Therefore 1 mole of Helium contains 6.022×10^{23} molecules.
  - Therefore 1 mole of Argon contains 6.022×10^{23} molecules.
  - Etc --- in other words it applies to "all the elements" of the periodic table

5. The Ideal Gas Law (Pressure, Volume, Temperature, and Molar Amount)

Standard Temperature and Pressure (STP)

- The reference condition for gases, chosen by convention to be **exactly 0°C and 1 atm pressure**.
- The molar volume, V_m, of a gas at STP is **22.4 L/mol.**

- These laws can be combined into **one law and one equation that represents them all.** This is called the **Ideal Gas Law** and it is represented by an equation.

Equation of the Ideal gas law is:

$$\boxed{PV = nRT}$$

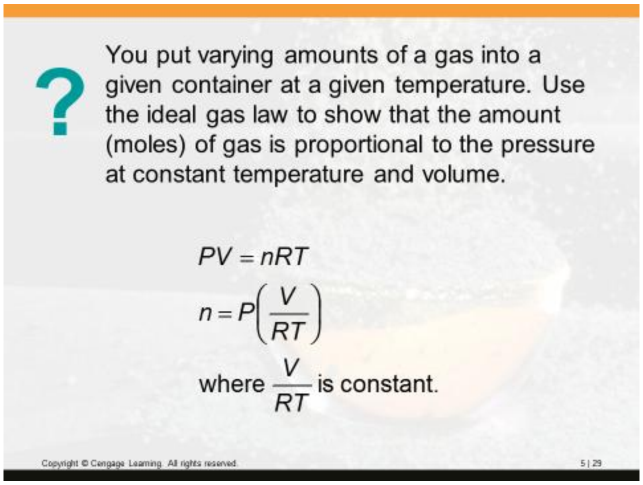
P = pressure *(Boyle's and Charles's Law)*
V = volume *(Boyle's and Charles's Law)*
T = temperature *(Boyles's and Charles's Law)*
n = number of moles of gas *(Avogadro's Law)*
R = molar gas constant (constant of proportionality relating the molar volume of a gas to temperature and pressure) *(Avogadro's Law)* ***R = 0.08206 L•atm /(K•mol)***

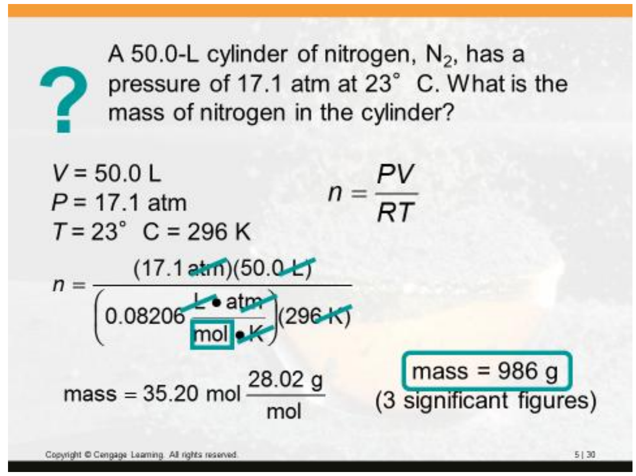
- The ideal gas law includes all information contained in Boyle's, Charles's and Avogadro's Law.
- In fact if you start with the ideal gas law you can determine any other gas law.

Table 5.5 Molar Gas Constant in Various Units

Value of *R*
0.082058 L·atm/(K·mol)
8.3145 J/(K·mol)*
8.3145 $kg \cdot m^2/(s^2 \cdot K \cdot mol)$
8.3145 $kPa \cdot dm^3/(K \cdot mol)$
1.9872 cal/(K·mol)*

Example – Calculations Using the Ideal Gas Law





Example – Calculations Using the Ideal Gas Law (p.145)

How many grams of oxygen (O_2), are there in a 50.0 L gas cylinder at 21°C when the oxygen pressure is 15.7 atm?

In asking for the mass of oxygen, we are in effect asking for moles of gas, *n*. the problem gives , *V*, and *T*, so you can use the ideal gas law to solve for *n*. The proper value to use for *R* depends on the units of *P* and *V* .

P = 15.7 atm
V = 50.0 L
T = (21 + 273) K = 294 K
n = ?

Solving the ideal gas law for *n* gives :

$n = PV/RT$

$n = 15.7\ \cancel{atm}\ \text{X}\ 50.0\ \cancel{L}\ /\ 0.0821\ \cancel{L}\bullet\cancel{atm}/(\cancel{K}\bullet \text{mol})\ \text{X}\ 294\ \cancel{K} = 32.5\ \text{mol}$

and converting moles to mass of oxygen yields:

$32.5\ \cancel{mol\ O_2}\ \text{X}\ 32.0\ g\ O_2 / 1\ \cancel{mol\ O_2} = \mathbf{1.04\ X\ 10^3\ g\ O_2}$

Example – Gas Density: Molecular-Weight Determination (p.146)

What is the density of oxygen (O_2), in grams per liter at 25°C and 0.850 atm?

The density of a gas is often expressed in g/L. Using the ideal gas law, you can calculate the moles of O_2 in 1 L of O_2 . Next you can convert the moles of O_2 to a mass of O_2 , keeping in mind that this mass is the amount of O_2 per liter of O_2 (g/l L), which is the density.

P = 0.850 atm
V = 1 L (exact value)
T = (25 + 273) K = 298 K
n = ?

Therefore:

$n = PV/RT$

$n = 0.850 \cancel{\text{atm}} \times 1 \cancel{\text{L}} / 0.0821 \cancel{\text{L}} \bullet \cancel{\text{atm}}/(\cancel{\text{K}} \bullet \text{mol}) \times 298 \cancel{\text{K}} = 0.0347 \text{ mol}$

Now converting mol O_2 to grams :

$0.0347 \cancel{\text{mol } O_2} \times 32.0 \text{ g } O_2 / 1 \cancel{\text{mol } O_2} = \mathbf{1.11 \text{ g } O_2}$

Therefore, the density of O_2 at $25^{o}C$ and 0.850 atm is **1.11 g/L** .

6. Dalton's Law of Partial Pressure

- While studying the composition of air, Dalton concluded that each gas in a mixture behaves as if it is the only gas present.

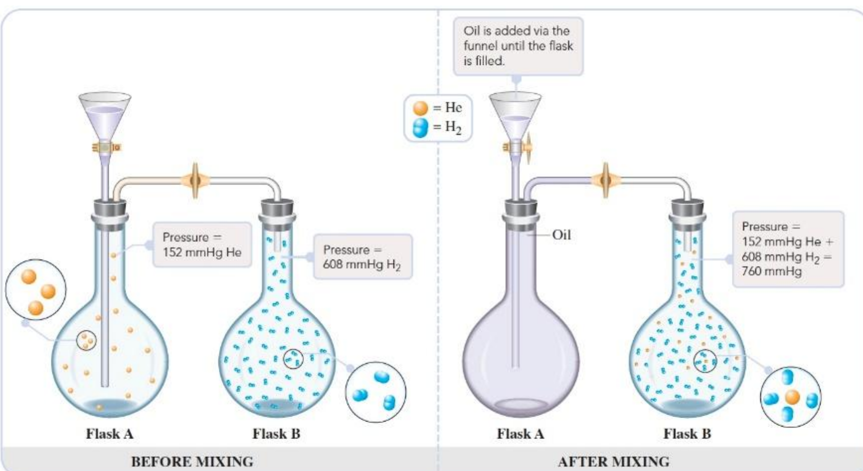
<u>Example</u>:

Flask 1 = A 1L flask is filled with hydrogen to the pressure of 608 mmHg.
Flask 2 = A 1L flask filled with helium to the pressure of 152 mmHg.
The helium if added to the flask with the hydrogen.

After the gases are mixed;

1) Each gas still occupies 1L the same as before.
2) Each gas has the same pressure.
3) Therefore the pressure exerted by the helium in the mixture is 152 mmHg and the pressure exerted by the hydrogen in the mixture is 602 mmHg.
4) The total pressure of the mixture, will add up to become:
   152 mmHg + 602 mmHg = 760 mmHg

(See Diagram p.149—Fig. 5.13 for the above example)



Originally (left), flask A contains He at 152 mmHg and flask B contains O_2 at 608 mmHg. Flask A is then filled with oil forcing the He into flask B (right). The new pressure in flask B is 760 mmHg.

- This is known as the partial pressure.
  - **Partial pressure =** *the pressure exerted by a particular gas in a mixture.*

Dalton's Law of partial pressure = *the sum of the partial pressures of all the different gases in a mixture is equal to the total pressure of the mixture.*

Example – Calculating Partial Pressures and Mole Fractions of a Gas in a Mixture (p.150)

A 1.00 L sample of dry air at 25^oC and 786 mmHg contains 0.925 g N_2 plus other gases, including oxygen, argon, and carbon dioxide.

a) What is the partial pressure (in mmHg) of N_2 in the air sample?
b) What are the mole fraction and mole percent of N_2 in the mixture?

a) Each gas in a mixture follows the ideal gas law. To calculate the partial pressure of N_2 , you convert 0.923 g N_2 to moles N_2 .

$0.923 \cancel{g\ N_2}$ X 1 mol N_2 / 28.0 $\cancel{g\ N_2}$ = 0.0330 mol N_2

You substitute into the ideal gas law (noting that 25^oC is 298 K)

$P_{N2} = n_{N2}RT / V$

$= 0.0330\ \cancel{mol}$ X $0.082\ \cancel{L}\ \bullet atm/(\cancel{K \bullet mol})$ X $298\ \cancel{K}$ / $1.00\ \cancel{L}$
= 0.807 atm **(=613 mmHg)**

b) The mole fraction of N_2 in air is .

Mole Fraction of N_2 = P_{N2} /P = 613 mmHg / 786 mmHg **= 0.780**

Air contains **78.0 mole percent** of N_2

Please note: Whenever you are solving for the mole fraction of a substance in a mixture, the number must always be less than 1. Therefore, if you obtain a value greater than 1, it is a sure sign that you have made an error.

YOUTUBE VIDEOS ON UNIT 8:

Boyle's Law:
https://www.youtube.com/watch?v=lE88aodUcTM

Charles Law:
https://www.youtube.com/watch?v=oIfFoiwRCVE

Avogadro's Law:
https://www.youtube.com/watch?v=r16dHmeqnsQ

Ideal Gas Law (Mr. Causey)
https://www.youtube.com/watch?v=6zpJOub2gDw

Law of Partial Pressure:
https://www.youtube.com/watch?v=RqffPYOoxd8

https://www.youtube.com/watch?v=gvXMg3J5j_Y