Gaseous State

Gases

- **The atmosphere that surrounds us is composed of a mixture of gases (air)**
- **The air we breath consists primarily of**
  - oxygen (about 21 %)
  - nitrogen (about 78 %)
- **Industrial and household uses of gases:**
- **To purify drinking water** chlorine
- **Welding** acetylene
- **Semiconductors** hydrogen chloride, hydrogen bromide

Gases

- **A substance that is normally in the gas state at ordinary temp. and pressure**
  - oxygen gas
- **Vapor: The gaseous form of any substance that is a liquid or solid at normal temp. and pressure**
  - water vapor or steam
- **The gas phase is the normal state for eleven of the elements**
  - H_2, N_2, O_2, and noble gases

Properties of Gases

- **Gases have no definite shape or volume and completely fill the container that holds them**
- **Gases can be compressed to a smaller volume or expanded to a larger volume**
- **The behavior of gases can be described by simple quantitative relationships: Gas Laws**

Properties of Gases

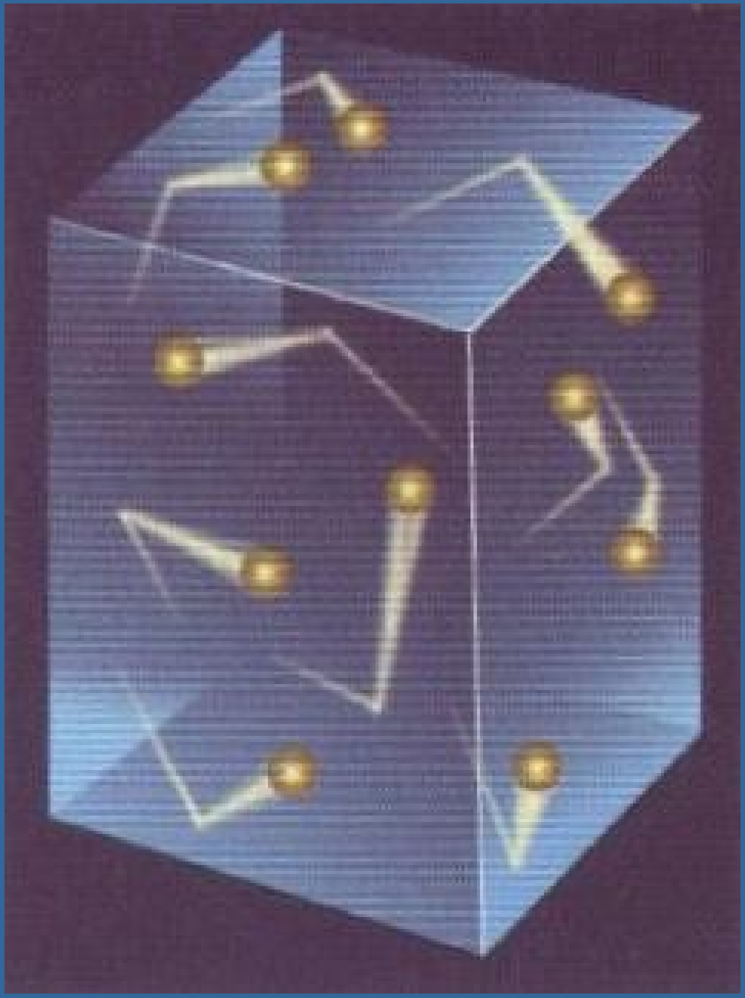
- **Only behavior of gases can be described by these simple quantitative relationships**
- **The gas laws are mathematical equations that predict how gases behave**
- **Why do they behave in a predictable manner?**
  - Why does a gas expand when its temperature is increased?
  - A model has been proposed to understand the physical properties of gases

Kinetic Molecular Theory

- **A mathematical model to explain the behavior of gases (the properties of an "ideal gas")**
- **The theory of moving molecules. It gives an understanding of pressure and temperature at a molecular level**
  - Gases are made up of tiny particles (atoms or molecules) widely separated from one another and in constant rapid motion
  - The particles have no attraction or repulsion for one another
  - The particles of a gas are extremely small compared to the total volume which contains them: A gas is mostly empty space

Kinetic Molecular Theory

- The average kinetic energy (motion) of the gas particles is proportional to the temperature of the gas (in Kelvin)
- Particles are in constant, random straight-line motion, colliding with each other and the container walls (causing pressure)



Properties of Gases

- **Gas laws are generalizations that describe mathematically the relationships of these four properties:**
  - –Pressure (P)
  - –Volume (V)
  - –Temperature (T)
  - –Amount of gas in moles (n)

Pressure

- A gas exerts pressure as its molecules collide with the surface of the container that holds it
- The pressure is the result of these collisions (impacts) divided by the unit area receiving the force

$$P = \frac{force}{area}$$

- The atmosphere exerts pressure on the earth by the collisions of molecules with every surface it contacts

Pressure

- **Caused by air being pulled towards earth by gravity**
- **Weather can change pressure**
  - Lows and Highs on the weather map
- **Altitude can change pressure**
  - Atmospheric pressure is lower in Denver than in Stockton because of “thinner air” due to less collisions of gas with surface

Volume

The volume of a gas depends on

1. The pressure of the gas
2. The temperature of the gas
3. The number of moles of the gas

The volume of a gas is expressed in mL and L

Temperature

- **Temperature of the gas is related to the KE of the molecule**
- **The velocity of gas particles increases as the temp. increases**

When working problems, temperature is expressed in Kelvin not Celcius

Amount of Gas

- Gases are usually measured in grams
- Gas law calculations require that the quantity of a gas is expressed in moles
- Moles of a gas represents the number of particles of a gas that are present

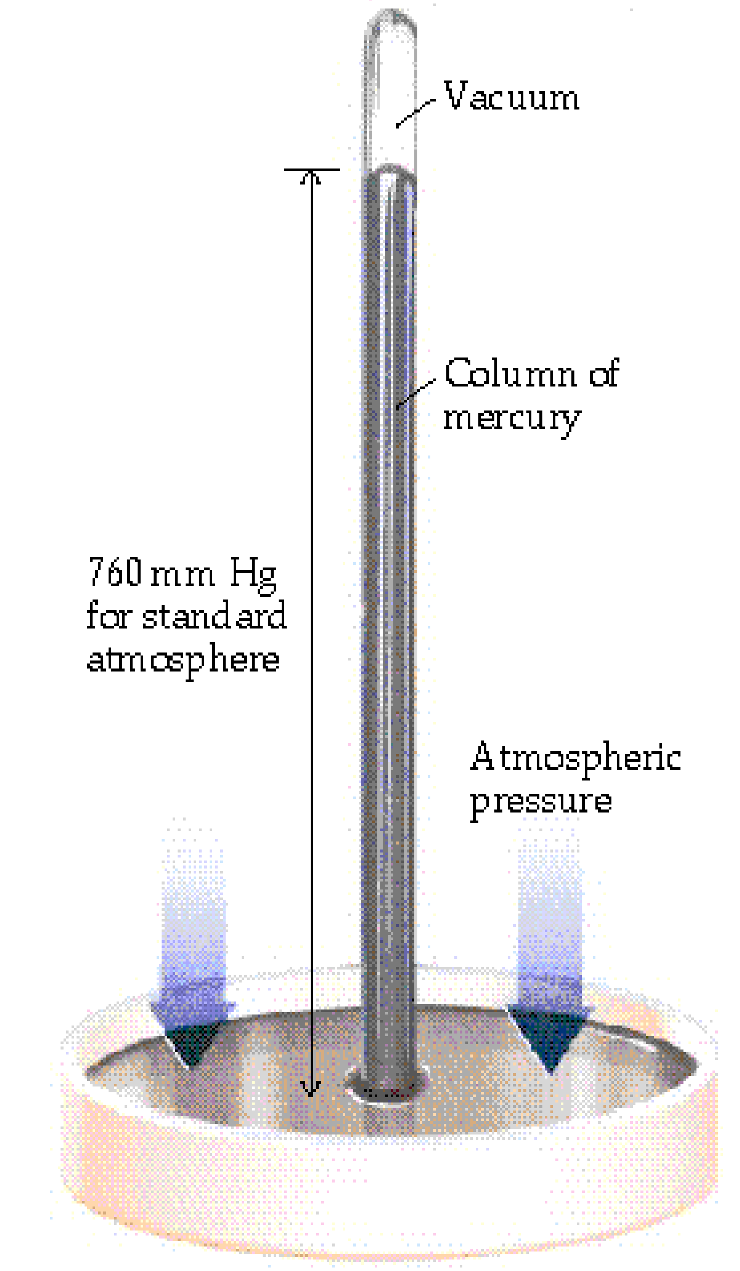
When working gas law problems,
convert mass (in grams) to moles

Gas Pressure

- **Use a device called a barometer**
  - Measures the height of Hg in a glass tube
  - Invented in 1643 by Torricelli
  - Units are in mm Hg
- **Common Units**
  - mm Hg or Torr (1 mm Hg = 1 Torr)
  - Atmospheres (atm) 1 atm = 760 Torr
  - Pascal (Pa) 1 atm = 101,325 Pa

Gas Pressure

- **Use a barometer**
  - Measure the height of a column of mercury (in mm)
  - The pressure of the atmosphere supports the column of mercury
  - The length of mercury in the column is a measure of the pressure
  - At sea level, the mercury would be 760 mm above the surface of the table



Pressure

- **Pressure depends on**
  - –The collisions of gas molecules with the walls of the container
  - –The **number** of collisions
  - –The **average force** of the collisions

Pressure and Volume

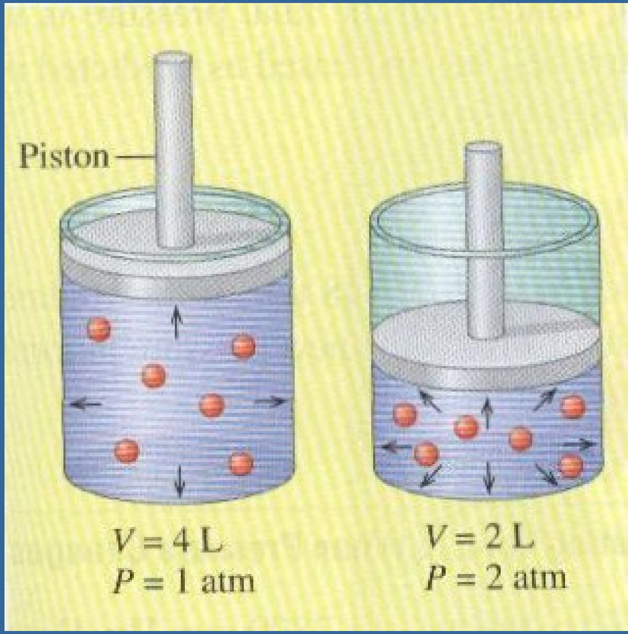
- **Boyle, R. discovered the relationship between gas pressure to gas volume**

If the pressure on the gas increases, then volume decreases proportionally

- **As long as *T* and *n* are constant**
- **As the volume decreases, more collisions with the walls occur (less space), so pressure increases**

Boyle's Law

- **Determined this from observations, not the theory**
- **Only pressure and volume are allowed to change**
- **Inverse relationship between P and V**
- **The temperature and the amount of gas (*n*) are held constant**



$$p \propto \frac{1}{V} \quad pV = k$$

$$P_1V_1 = P_2V_2$$

Temperature and Volume

- Charles, J. found that the volume of any gas will increase with an increase in temperature
- The volume of a sample of gas is directly proportional to its temperature (K)
- As long as P and n are held constant

Charles's Law

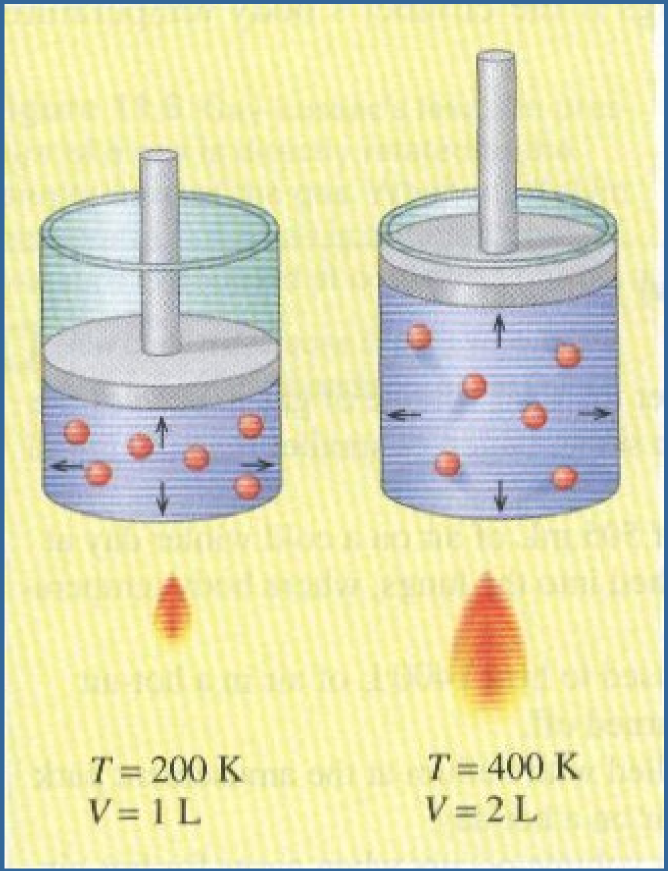
- **What happens when temperature increases?**
  - Increase speed of molecules
  - Increase force with which they hit the walls
- **If volume can change**
  - Force "stretches" out its container until it ends up with the same pressure it had before, just a larger volume

Charles's Law

- **Volume and temperature are proportional (if pressure and amount of gas is constant)**

$$V \propto T$$

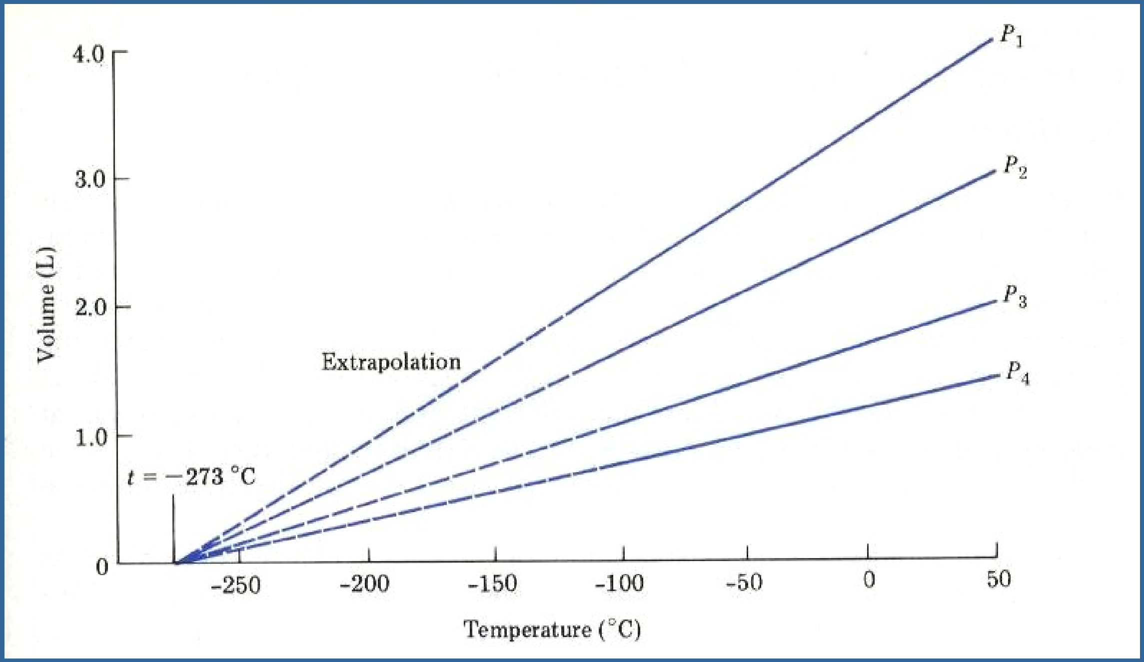
  - Increase temperature, volume increases
  - Decrease temperature, volume decreases



$$V = bT \qquad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Charles's Law

- **If we could cool down a gas enough, its volume would reduce to zero**
  - Absolute Zero (0 Kelvin)
  - Theoretical only because at a low enough temperature, a gas will liquify



Temperature and Pressure

- Gay-Lussac, J. L. discovered the relationship between gas pressure and temperature of a gas

The pressure of a sample of gas, at a constant volume, is directly proportional to its temperature (K)

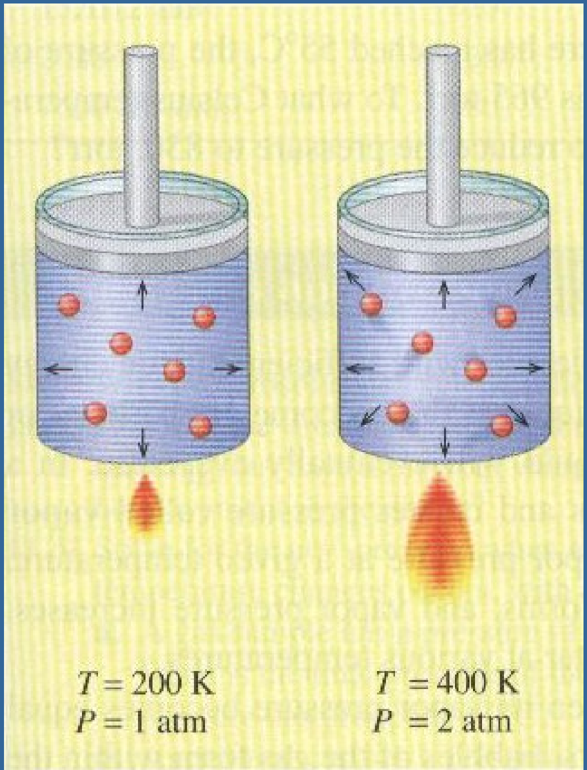
- As long as V and *n* are constant

$$P \propto T$$

- By KM theory, as the temperature increases, the speed of the gas molecules increases

Temperature and Pressure

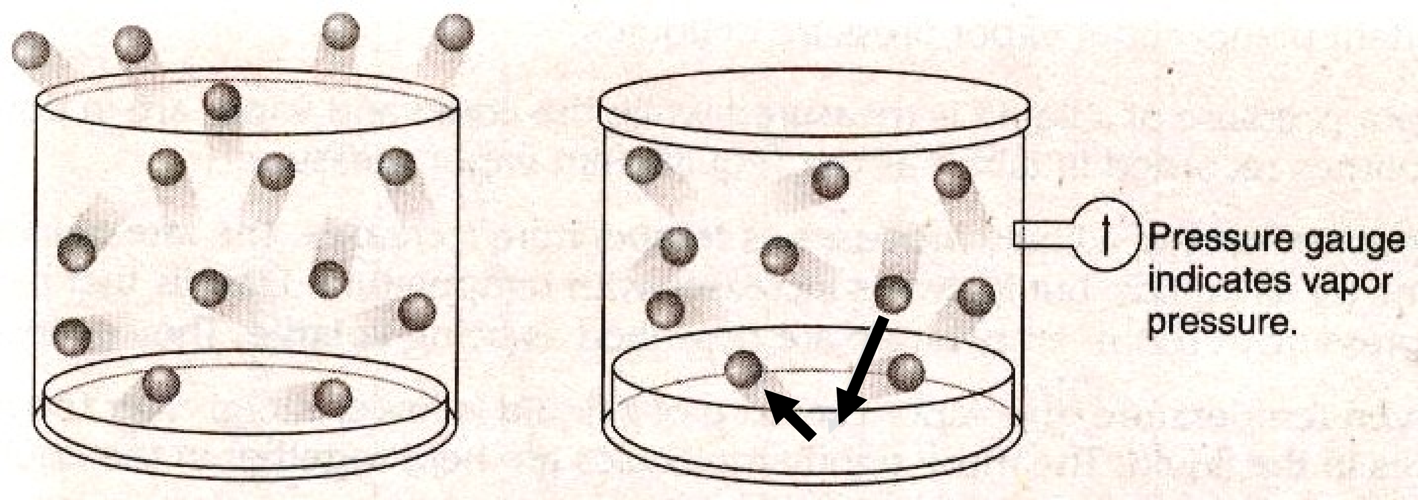
- **By KM theory, as the temperature increases, the speed of the gas molecules increases**
- **Increase in the number of collisions the molecules have with the container wall**
- **Thus, an increase in**



$$\frac{P}{T} = k \qquad \frac{\mathrm{P}_1}{\mathrm{T}_1} = \frac{\mathrm{P}_2}{\mathrm{T}_2}$$

Vapor pressure and Boiling Point

- **Liquids are constantly evaporating**
- **Molecules at the surface of the liquid will break away from the liquid and enter into the gas phase**
- **But in a closed container the molecules cannot escape to the atmosphere**



Vapor pressure and Boiling Point

- **At a certain point molecules will continue to evaporate and at the same rate molecules will return to the liquid**
- **At equilibrium the rates of vaporization and condensation are equal**

Liquid to Gas = Gas to Liquid

- **Vapor Pressure of a liquid is the pressure exerted by its vapor at equilibrium**

Vapor pressure and Boiling Point

- The vapor pressure of a liquid will increase with increase in temperature
- The normal **boiling point** of a liquid is the temperature at which its vapor pressure equals the atmospheric pressure
- An increase in altitude will lower the atmospheric pressure so liquids will boil at a lower temperature

The Combined Gas Law

- **An expression which combines Boyle's, Charles's, and Gay-Lussac's Law**

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

- **Applies to all gases and mixtures of gases**
- **If five of the six variables are known, the sixth can be calculated**

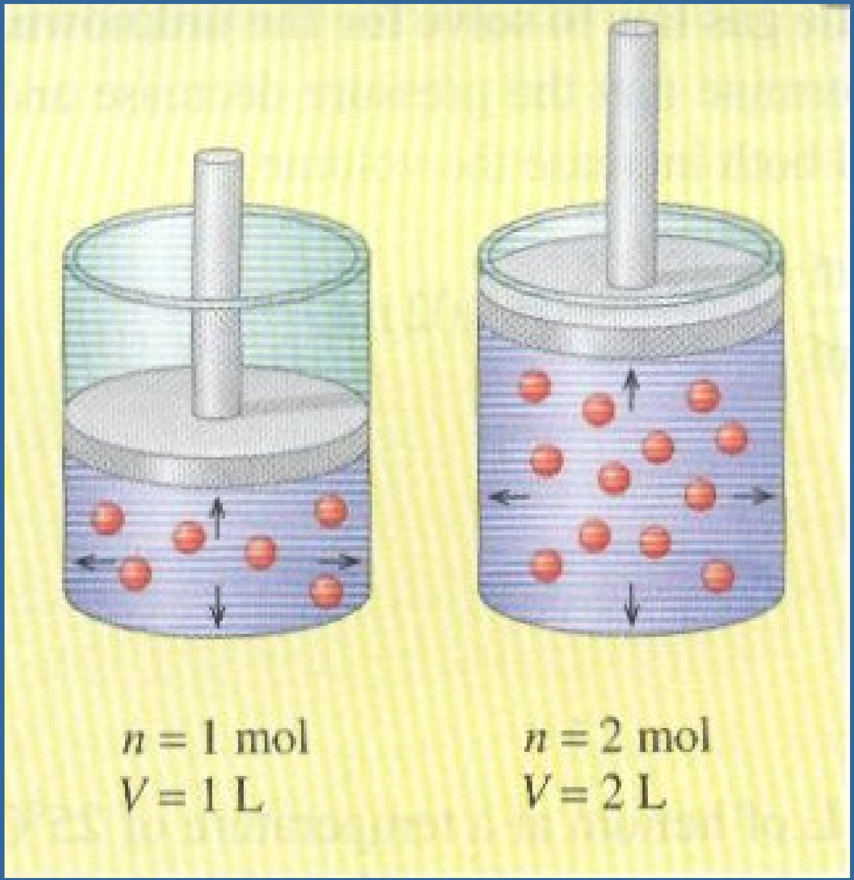
Volumes and Moles

- **Avogadro proposed the relationship between volume and a quantity of gas (in moles, n)**
- **Equal volumes of two different gases at the same temperature and pressure will contain the same number of molecules**

The volume of a gas varies directly with the number of moles of gas as long as the pressure and temperature remain constant

Avogadro’s Law

- **If you increase (double) the number of moles, you increase (double) the volume**
- **If the number of molecules is cut in half, the volume will be cut in half**
  - More molecules push on each other and the walls, “stretching” the container and making it larger until the pressure is the same as before



$$V \propto n \qquad \frac{V_1}{n_1} = \frac{V_2}{n_2}$$

$$V = an$$

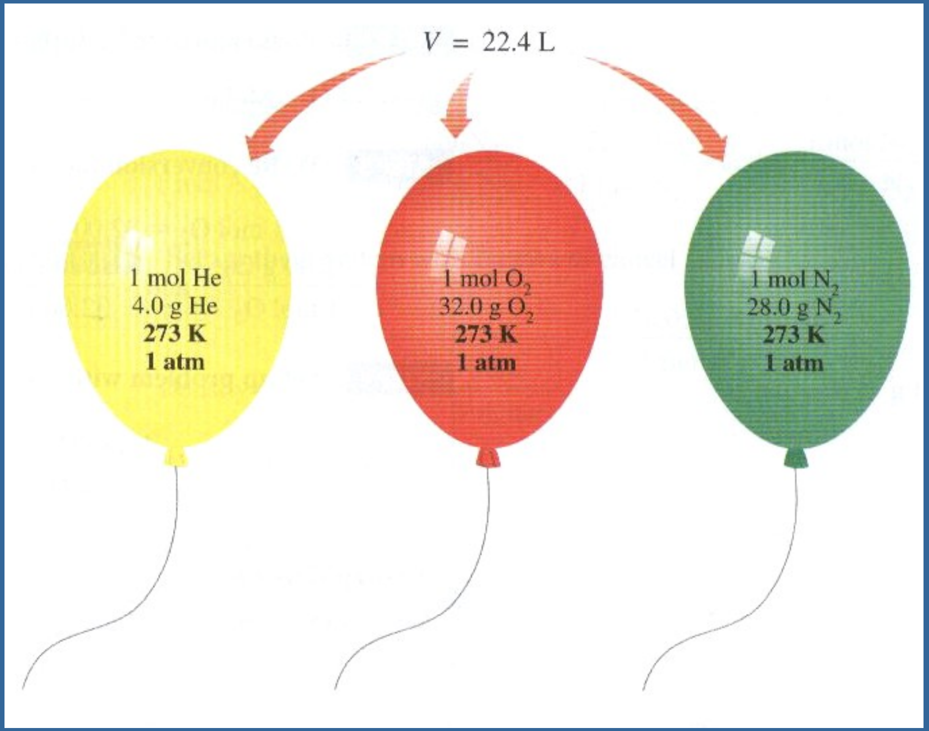
STP and Molar Volume

- **Gas volumes can only be compared at the same T and P**
- **The universally accepted conditions are T = 0 °C and P = 1.00 atm**

Standard Temperature and Pressure = STP
Standard Temperature = 0 °C (273 K)
and Standard Pressure = 1.00 atm (760 mmHg)

STP and Molar Volume

- **The volume of one mole of <u>any</u> <u>gas</u> at STP is 22.4 L**
  - Known as the standard molar volume
  - When used in calculations at STP, the conversion factor(s) can be used:



$$\frac{1\text{ mol gas}}{22.4\text{ L}} = \frac{22.4\text{ L}}{1\text{ mol gas}}$$

Density of a Gas at STP

At the same T and P, one mole of any gas will occupy the same volume

- **At STP, the density of a gas depends on its molar mass**
- **The higher the molar mass, the greater the density**
  - CO_2 has a higher molar mass than N_2 and O_2, so it is more dense than air
  - CO_2 released from a fire extinguisher will cover the fire and prevent O_2 from reaching the combustible material
  - Helium is less dense than air so a balloon filled with helium rises in the air

The Ideal Gas Law

- **The ideal gas law describes the relationships among the four variables: P, T, V, n (moles)**
- **Derived from the three different relationships concerning the volume of a gas**

$$V = k \times \frac{1}{P}$$

Boyle's Law
n, T constant

$$V = kt$$

Charles's Law
n, P constant

$$V = kn$$

Avogadro's Law
P, T constant

$$V = \frac{\boxed{R}\, t\, n}{P}$$

Combined into a single expression

The Ideal Gas Law

- The ideal gas law describes the relationships among the four variables:

P, T, V, n (moles)

- Derived from the three different relationships concerning the volume of a gas

Ideal Gas Law

- **The usual written form where all three gas laws are put together into one equation**
  - Combines Boyle’s Law, Charles’s Law and Avogadro’s Law into one relationship

$$pV = nRT$$

Ideal Gas Law

- **R is the gas law constant**
  - Make sure that you use the correct units in the ideal gas law with the correct version of R

$$R = 8.314 \frac{J}{mol \bullet K} \text{ or } 0.0821 \frac{L \bullet atm}{mol \bullet K}$$

Ideal Gas Law

- **Ideal gas law can be re-arranged to make the equations from Boyle's, Charles's or Avogadro's Laws**
- **Boyle's Law**
  - Temperature and number of moles are constant

$$p_1 V_1 = nRT = p_2 V_2$$

$$p_1 V_1 = p_2 V_2$$

Ideal Gas Law

- **Charles's Law**
  - Pressure and number of moles are constant

$$\frac{V_1}{T_1} = \frac{nR}{p} = \frac{V_2}{T_2} \qquad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

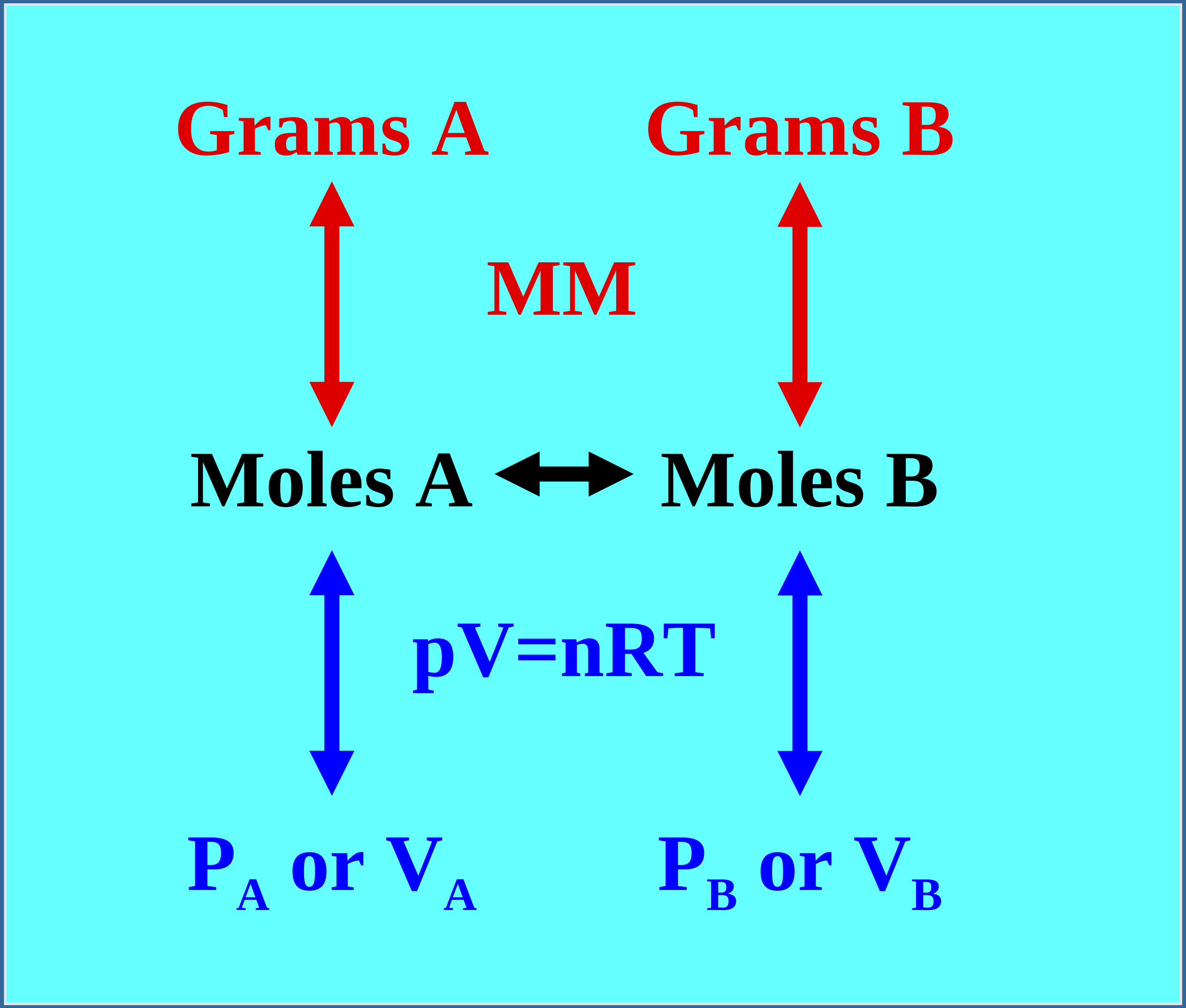
- **Avogadro's Law**
  - Pressure and Temperature are constant

$$\frac{V_1}{n_1} = \frac{TR}{p} = \frac{V_2}{n_2} \qquad \frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Using the Ideal Gas Law

- **Allow the direct calculation of P, T, V, or n**
- **If three of the four variables are known, (i.e. P, T, n) the unknown variable, V can be calculated**
- **Allow the calculation of the molar mass of a gas**
- **Can use in stoichiometric calculations with chemical reactions**

Using the Ideal Gas Law



Using the Ideal Gas Law

- The ideal gas equation is used when three of the four variables are known
- Problems are given with one set of variables given and one missing variable
- Solve Ideal Gas Law by rearranging the equation so that the missing quantity is on one side of the equation and everything that stays constant is on the other

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

Using the Combined Gas Law

- Allows you to calculate a change in one of the three variables (P, V, T) caused by a change in both of the other two variables

1) Determine the initial values
2) Determine the final values
3) Set two equations equal to one another, one for initial values, one for final values
4) Solve for the unknown variable

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

Gas Laws and Chemical Reactions

- In chapter 8 we used the balanced equation and mass of one reactant to calculate the mass of any other component
- Mass to mass or gram to gram
- Can also do **volume to mass** and **mass to volume** problems where at least one of the components is a gas
- The ideal gas law will relate the moles of a gas to its P, T and V

Gas Laws and Chemical Reactions

- In chapter 8 we used the balanced equation and mass of one reactant to calculate the mass of any other component

Mole to mole or gram to gram problems

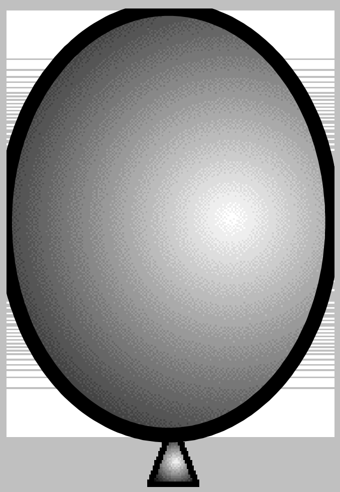
- In this chapter, where at least one of the components is a gas

Volume to mass and mass to volume problems

- The ideal gas law will relate the moles of a gas to its P, T and V

Ideal Gas Law Example 1

- A balloon is filled with 1300 mol of H_2. If the temperature of the gas is 23 °C and the pressure is 750 mm Hg, what is the volume of the balloon?



1300 mol H_2

23 °C

750 mm Hg

V = ?

Ideal Gas Law Example 1

$$p = \frac{750\ \cancel{\text{torr}}}{} \left| \frac{1\ \text{atm}}{760\ \cancel{\text{torr}}} \right. = 0.987\ \text{atm}$$

$$n = 1300\ \text{mol}$$

$$T = 23 + 273 = 296\ \text{K}$$

$$R = 0.082\ \frac{\text{L}\bullet\text{atm}}{\text{mol}\bullet\text{K}}$$

$$V = ?$$

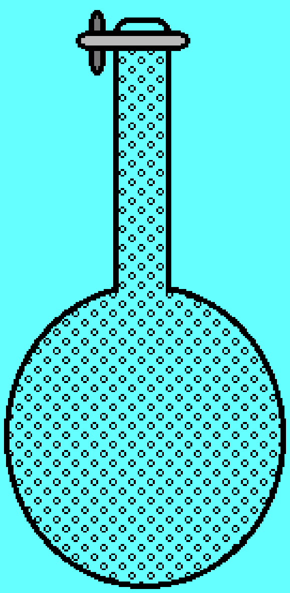
$$V = \frac{nRT}{p}$$

$$V = \frac{(1300\ \cancel{\text{mol}})(0.082\ \frac{\text{L}\bullet\cancel{\text{atm}}}{\cancel{\text{mol}}\bullet\cancel{\text{K}}})(296\ \cancel{\text{K}})}{(0.987\ \cancel{\text{atm}})}$$

$$V = 31{,}969.2\ \text{L}$$

Ideal Gas Law Example 2
Determining the Molar Mass of a Gas

- **A 0.105 g sample of gaseous compound has a pressure of 561 mm Hg in a volume of 125 mL at 23.0 °C. What is its molar mass?**



$$p = \frac{561\cancel{\text{torr}}}{} \left| \frac{1\text{atm}}{760\cancel{\text{torr}}} \right. = 0.738 \text{ atm}$$

$$T = 23 + 273 = 296 \text{ K}$$

$$V = 125 \text{ mL} = 0.125 \text{ L}$$

$$g = 0.105 \text{ g}$$

Ideal Gas Law Example 2

$$n = \frac{pV}{RT}$$

$$n = \frac{(0.78\,\cancel{\text{atm}})(0.125\,\cancel{\text{L}})}{\left(0.0821\,\frac{\cancel{\text{L}}\cdot\cancel{\text{atm}}}{\text{mol}\cdot\cancel{\text{K}}}\right)(296\,\cancel{\text{K}})}$$

$$n = 4.004\times10^{-3}\ \text{mol}$$

$$MW = \frac{g}{\text{mol}} = \frac{0.105\ \text{g}}{4.004\times10^{-3}\ \text{mol}}$$

$$\boxed{MW = 26.2\ \text{g/mol}}$$

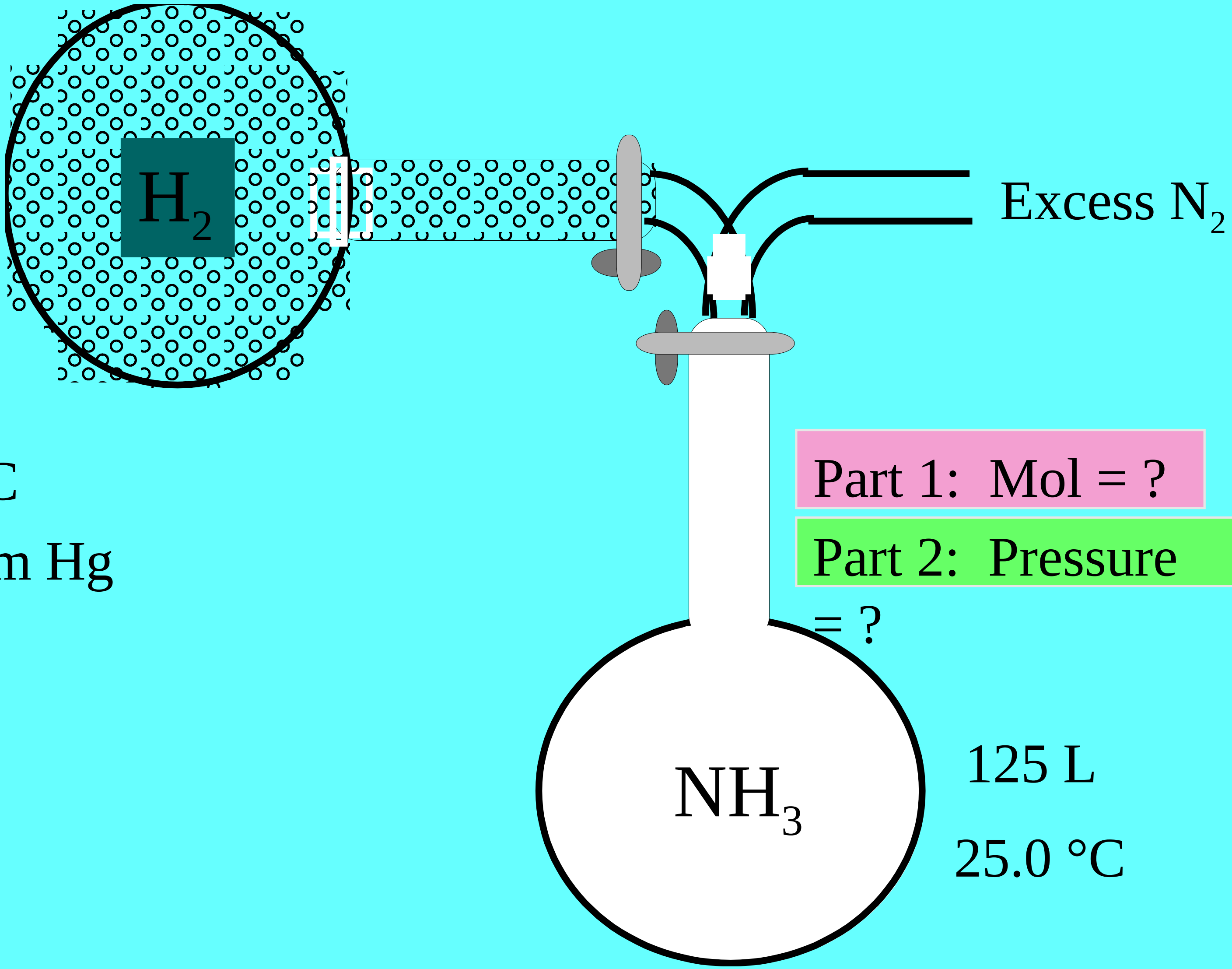
Ideal Gas Law Example 3

- Gaseous ammonia is synthesized by the following reaction.

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

- Suppose you take 355 L of H_2 gas at 25.0 °C and 542 mm Hg and combine it with excess N_2 gas. What quantity of NH_3 gas, in moles, is produced? If this amount of NH_3 gas is stored in a 125 L tank at 25 °C, what is the

Ideal Gas Law Example 3



355 L

25.0 °C

542 mm Hg

Part 1: Mol = ?

Part 2: Pressure = ?

125 L

25.0 °C

Ideal Gas Law Example 3

1) Find the moles of hydrogen using ideal gas law

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

$$p = \frac{542\ \cancel{\text{torr}}}{} \frac{1\ \text{atm}}{760\ \cancel{\text{torr}}} = 0.713\ \text{atm}$$

$$V = 355\ \text{L}$$

$$T = 25 + 273 = 298\ \text{K}$$

$$n = ?$$

$$n = \frac{pV}{RT}$$

$$n = \frac{(0.713\ \cancel{\text{atm}})(355\ \cancel{\text{L}})}{(0.0802\ \frac{\cancel{\text{L}}\bullet\cancel{\text{atm}}}{\text{mol}\bullet\cancel{\text{K}}})(298\ \cancel{\text{K}})}$$

$$n = 10.59\ \text{mol}\ H_2$$

Ideal Gas Law Example 3

2) Determine the moles of ammonia produced using the mole-mole factor in the balanced equation

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

3 moles H_2 = 2 moles NH_3

$$\frac{2\ \text{mol}\ NH_3}{3\ \text{mol}\ H_2}\ \textit{and}\ \frac{3\ \text{mol}\ H_2}{2\ \text{mol}\ NH_3}$$

Now, combine steps 1 and 2 to calculate moles

$$\frac{10.59\ \cancel{\text{mol}\ H_2}}{} \left| \frac{2\ \text{mol}\ NH_3}{3\ \cancel{\text{mol}\ H_2}} \right. = \boxed{7.06\ \text{mol}\ NH_3}$$

Ideal Gas Law Example 3

3) Calculate the pressure of the ammonia gas

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

$$T = 25 + 273 = 298\ K$$

$$n = 7.06\ mol$$

$$V = 125\ L$$

$$p = ?$$

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

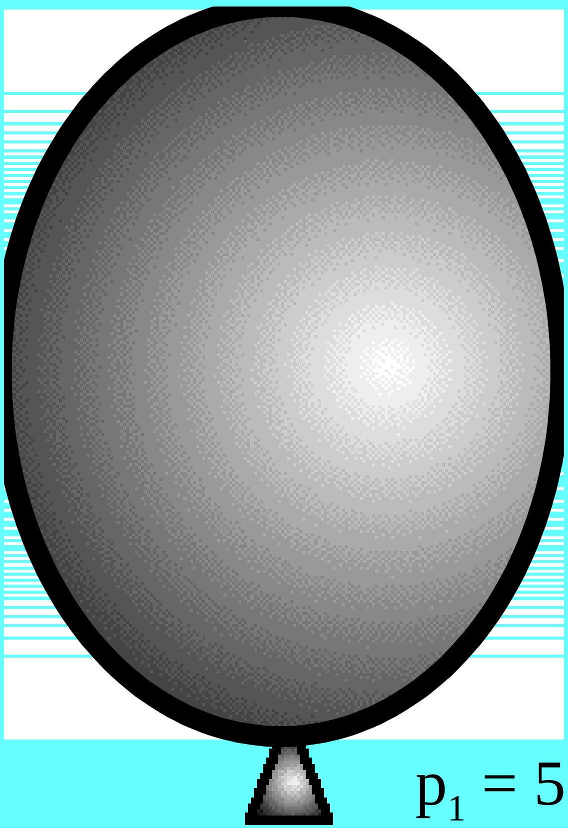
$$p = \frac{(7.06\ mol)(0.082)(298\ K)}{(125\ L)}$$

$$p = 1.38\ atm$$

Ideal Gas Law Example 4
Using Boyle’s Law

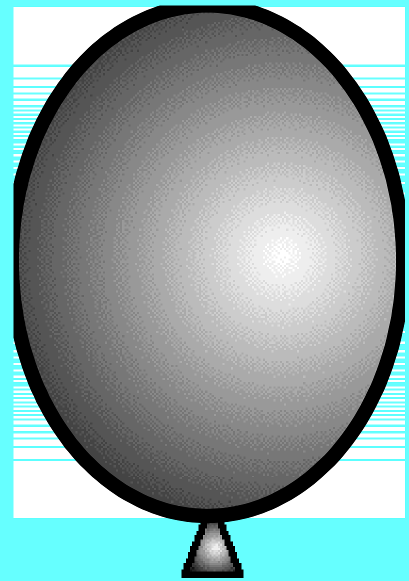
- **A sample of CO_2 has a pressure of 55 mm Hg in a volume of 125 mL. The sample is compressed so that the new pressure of the gas is 78 mm Hg. What is the new volume of the gas? (The temperature does not change in this process.)**

Ideal Gas Law Example 4



$p_1 = 55$ mm Hg

$V_1 = 125$ mL



$p_2 = 78$ mm Hg

$V_2 = ?$

Ideal Gas Law Example 4

$$p_1 = 55 \text{ mm Hg}$$

$$V_1 = 125 \text{ ml}$$

$$p_2 = 78 \text{ mm Hg}$$

$$V_2 = ?$$

$$pV = nRT$$

$$p_1 V_1 = p_2 V_2$$

$$\frac{p_1 V_1}{p_2} = V_2$$

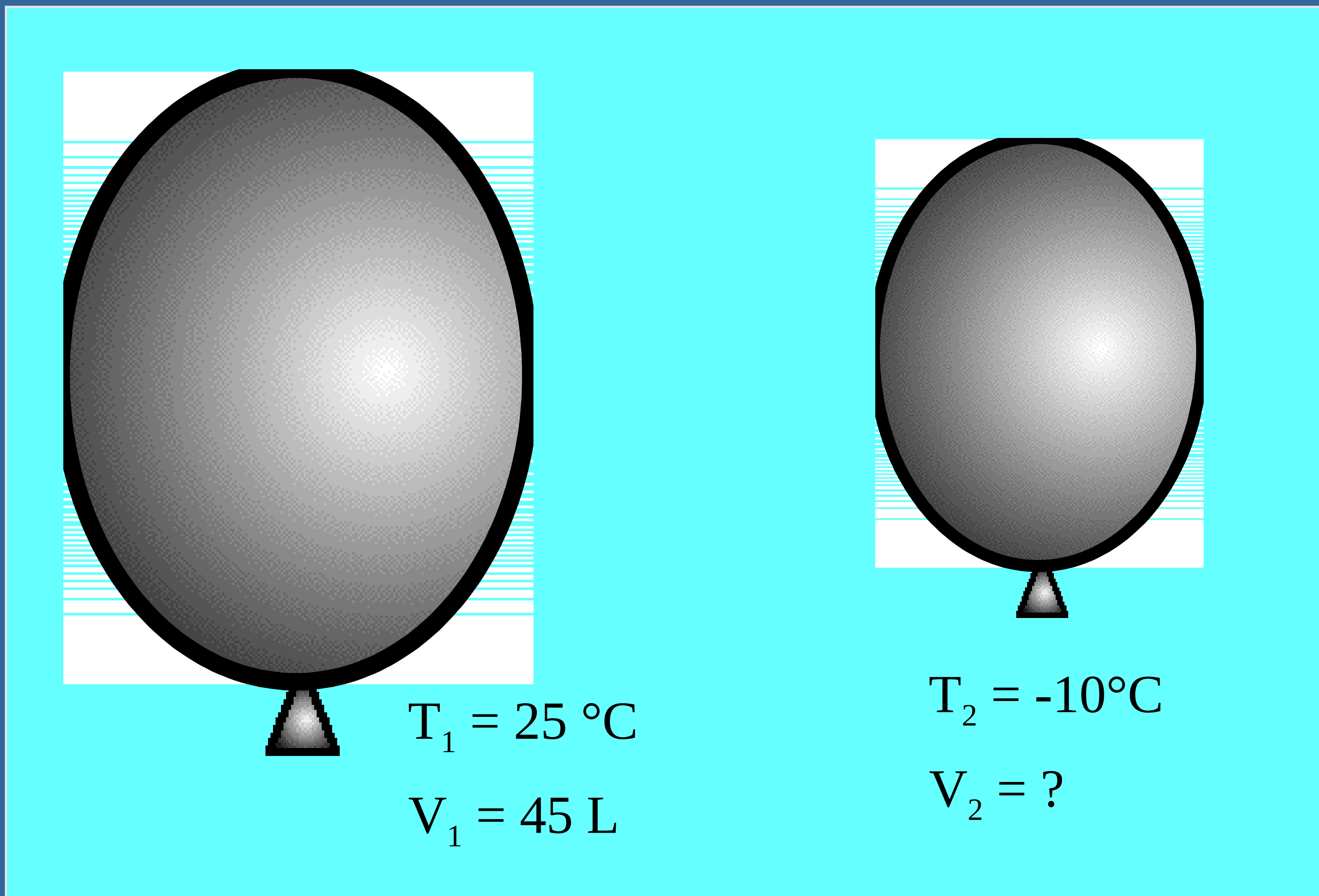
$$V_2 = \frac{(55 \cancel{\text{ mm Hg}})(125 \text{ ml})}{(78 \cancel{\text{ mm Hg}})}$$

$$V_2 = 88.14 \text{ ml}$$

Ideal Gas Law Example 5
Using Charles's Law

- **A balloon is inflated with helium to a volume of 45 L at room temperature (25 °C). If the balloon is cooled to -10 °C, what is the new volume of the balloon? (Assume that the pressure does not change.)**

Ideal Gas Law Example 5



Ideal Gas Law Example 5

$$T_1 = 25 + 273 = 298\ \text{K}$$

$$V_1 = 45\ \text{L}$$

$$T_2 = -10 + 273 = 263\ \text{K}$$

$$V_2 = ?$$

$$pV = nRT$$

$$\frac{V}{T} = \frac{nR}{p}$$

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

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

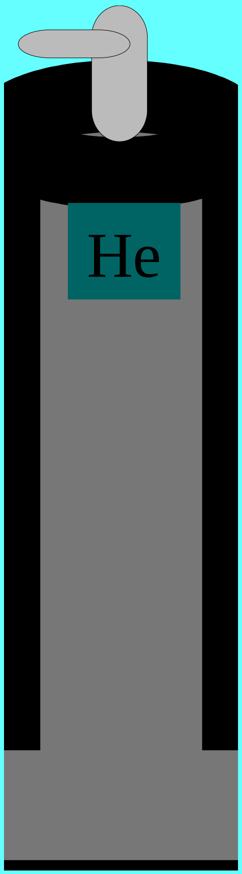
$$V_2 = \frac{(263\ \cancel{\text{K}})(45\ \text{L})}{(298\ \cancel{\text{K}})}$$

$$V_2 = \boxed{40.\ \text{L}}$$

Ideal Gas Law Example 6

- **You have a 22 L cylinder of helium at a pressure of 150 atm and at 31 °C. How many balloons can you fill, each with a volume of 5.0 L, on a day when the atmospheric pressure is 755 mm Hg and the temperature is 22 °C?**

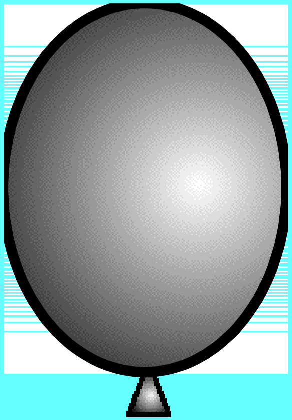
Ideal Gas Law Example 6



$V_1 = 22$ L

$p_1 = 150$ atm

$T_1 = 31$ °C



$V_2 = ?$

$p_2 = 755$ mm Hg

$T_2 = 22$ °C

V/balloon = 5 L/balloon

Ideal Gas Law Example 6

$$p_1 = 150 \text{ atm}$$

$$T_1 = 31 + 273 = 304 \text{ K}$$

$$V_1 = 22 \text{ L}$$

$$p_2 = \frac{755 \cancel{\text{torr}}}{} \left| \frac{1 \text{ atm}}{760 \cancel{\text{torr}}} \right. = 0.99 \text{ atm}$$

$$T_2 = 22 + 273 = 295 \text{ K}$$

$$V_2 = ?$$

$$pV = nRT$$

$$\frac{pV}{T} = nR$$

$$\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}$$

$$\frac{p_1 V_1 T_2}{p_2 T_1} = V_2$$

Ideal Gas Law Example 6

$$V_2 = \frac{(150\ \cancel{atm})(22\ L)(295\ \cancel{K})}{(0.99\ \cancel{atm})(304\ \cancel{K})}$$

$$V_2 = 3234.6\ L$$

$$\frac{3234.6\ \cancel{L}}{} \left| \frac{1\ \text{balloon}}{5\ \cancel{L}} \right. = \boxed{646\ \text{balloons}}$$

Gas Mixtures

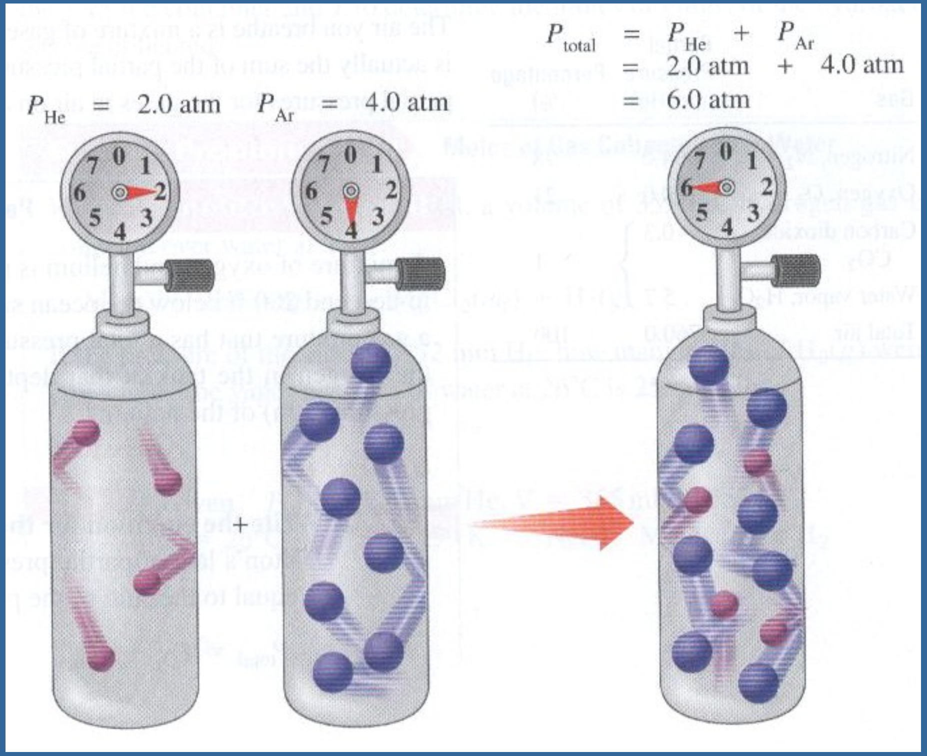
- **Air we breathe is a mixture of nitrogen, oxygen, carbon dioxide, water vapor and other gases**
  - Each gas exerts a pressure
  - Total of all of the pressures is the atmospheric pressure
- **In a mixture of gases in a container**
  - Each molecule moves in the container independently of the others
  - Each collides with the container as frequently as if it were the only gas present

Partial Pressures
Dalton's Law

- **Partial Pressure**
  - Pressure exerted by an individual gas in a mixture of gases
  - The pressure it would exert if were alone under the same conditions
- **Dalton's Law: The total pressure for a mixture of gases is the sum of partial pressures of the individual gases**

Dalton's Law

- Each gas behaves independently of any others in a mixture of ideal gases
- P_{total} is the total pressure, exerted by the gases in a mixture. It is directly related to n_{total}
- n_{total} is the sum of the moles of gases in the mixture



Dalton's Law

- **Can come up with an equation for P_{total} for a mixture of gases**

$$p_A = \frac{n_A RT}{V} \qquad p_B = \frac{n_B RT}{V} \qquad p_C = \frac{n_C RT}{V}$$

$$p_{tot} = p_A + p_B + p_C$$

$$p_{tot} = n_A \frac{RT}{V} + n_B \frac{RT}{V} + n_C \frac{RT}{V}$$

$$p_{tot} = n_{tot} \frac{RT}{V}$$

Partial Pressure Example 1

- **A halothane-oxygen mixture used for anesthesia is placed in a 5.0 L tank at 25.0 °C. You used 15.0 g of halothane vapor and 23.5 g of oxygen. What is the total pressure (in torr) of the gas mixture in the tank? What are the partial pressures (in torr) of the gases? (MW Halothane = 197.38g/mol)**

Partial Pressure Example 1

- **To solve:**
  - Find the total number of moles and solve for P_{tot}

$$\frac{15.0\ \cancel{\text{g H}}}{} \left| \frac{1\ \text{mol H}}{197.38\ \cancel{\text{g H}}} \right. = 0.076\ \text{mol H}$$

$$\frac{23.5\ \cancel{\text{g O}_2}}{} \left| \frac{1\ \text{mol O}_2}{32.0\ \cancel{\text{g O}_2}} \right. = 0.734\ \text{mol O}_2$$

Partial Pressure Example 1

$$n = 0.076 + 0.734 = 0.810 \text{ mol}$$

$$T = 25 + 273 = 298 \text{ K}$$

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

$$V = 5.0 \text{ L}$$

$$p = ?$$

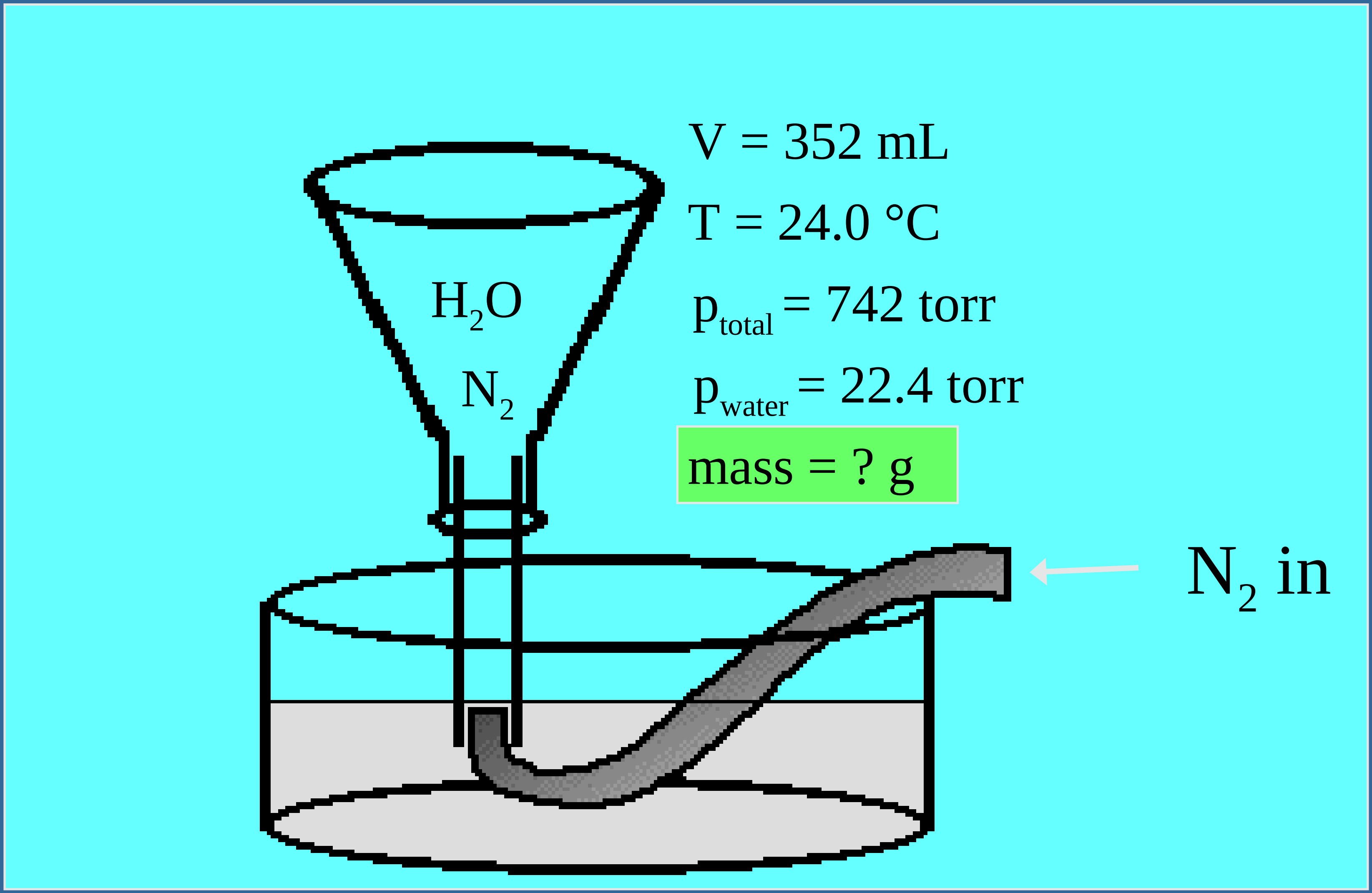
$$p = \frac{(0.810 \text{ mol})(0.082)(298 \text{ K})}{(5.0 \text{ L})}$$

$$p_{tot} = \frac{3.96 \cancel{\text{atm}}}{} \left| \frac{760 \text{ torr}}{1 \cancel{\text{atm}}} \right. = 3010 \text{ torr}$$

Partial Pressure Example 2

- **In an experiment, 352 mL of gaseous nitrogen is collected in a flask over water at a temperature of 24.0 °C. The total pressure of the gases in the flask is 742 torr. What mass of N_2 is collected? (Vapor pressure of water is 22.4 torr)**

Partial Pressure Example 2



Partial Pressure Example 2

$$p_{N_2} = 742 - 22.4 = 719.6 \text{ torr}$$

$$\frac{719.6\ \cancel{\text{torr}}}{} \left| \frac{1 \text{ atm}}{760\ \cancel{\text{torr}}} \right. = 0.947 \text{ atm}$$

$$n = \frac{pV}{RT}$$

$$V = 352 \text{ mL} = 0.352 \text{ L}$$

$$T = 24 + 273 = 297 \text{ K}$$

$$n = ?$$

$$n = \frac{(0.947 \text{ atm})(0.352 \text{ L})}{(0.082)(297 \text{ K})}$$

$$n = 0.0137 \text{ mol } N_2$$

$$\frac{0.0137\ \cancel{\text{mol } N_2}}{} \left| \frac{28.01 \text{ g } N_2}{1\ \cancel{\text{mol } N_2}} \right. = \boxed{0.384 \text{ g } N_2}$$

- end