Chapter 5

The Gas Laws

Pressure

- Force per unit area.
- Gas molecules fill container.
- Molecules move around and hit sides.
- Collisions are the force.
- Container has the area.
- Measured with a barometer.

Units of pressure

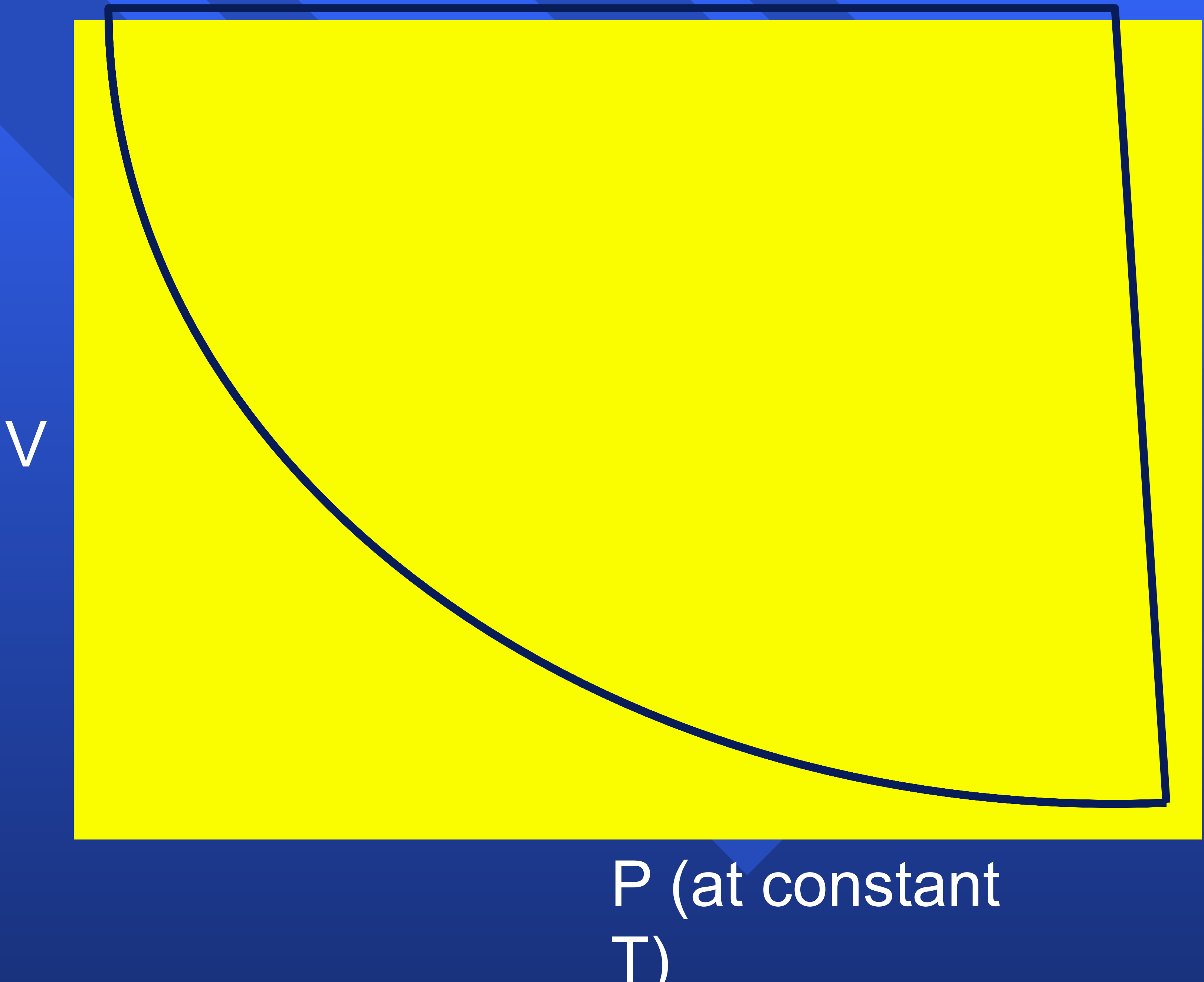
- 1 atmosphere = 760 mm Hg
- 1 atm = 760 Torr
- 1 atm = 101,235 Pascals = 101.325 kPa
- Can make conversion factors from these.
- What is 724 mm Hg in kPa?

Temperature: A measure of how much energy the particles in a gas have. Units of temperature that you'll run into include degrees Celsius and Kelvins (which is equal to 273 plus the degrees Celsius).

Volume: The amount of space that some object occupies.

The Gas Laws

- Boyle's Law
- Pressure and volume are inversely related at constant temperature.
- $PV = k$
- As one goes up, the other goes down.
- $P_1V_1 = P_2 V_2$
- Graphically


V
P (at constant T)

Examples

- **20.5 L of nitrogen at 25ºC and 742 torr are compressed to 9.8 atm at constant T. What is the new volume?**
- **30.6 mL of carbon dioxide at 740 torr is expanded at constant temperature to 750 mL. What is the final pressure in kPa?**

Charle’s Law

- **Volume of a gas varies directly with the absolute temperature at constant pressure.**
- **$V = kT$ (if T is in Kelvin)**
- **$V_1 = V_2$ $T_1 = T_2$**
- **Graphically**

Examples

- **What would the final volume be if 247 mL of gas at 22ºC is heated to 98ºC , if the pressure is held constant?**

Examples

- **At what temperature would 40.5 L of gas at 23.4ºC have a volume of 81.0 L at constant pressure?**

Avogadro's Law

- **Avagadro’s**
- **At constant temperature and pressure, the volume of gas is directly related to the number of moles.**
- **V = k n (n is the number of moles)**
- $V_1 = V_2 \qquad n_1 = n_2$

Gay- Lussac Law

- At constant volume, pressure and absolute temperature are directly related.
- $P = k\,T$
- $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Combined Gas Law

- **If the moles of gas remains constant, use this formula and cancel out the other things that don't change.**
- $$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Examples

- **A deodorant can has a volume of 175 mL and a pressure of 3.8 atm at 22ºC. What would the pressure be if the can was heated to 100.ºC?**
- **What volume of gas could the can release at 22ºC and 743 torr?**

Ideal Gas Law

- PV = nRT
- V = 22.41 L at 1 atm, 0ºC, n = 1 mole, what is R?
- R is the ideal gas constant.
- R = 0.0821 L atm/ mol K
- Tells you about how a gas is NOW.
- The other laws tell you about a gas when it changes.

Examples

- **A 47.3 L container containing 1.62 mol of He is heated until the pressure reaches 1.85 atm. What is the temperature?**

Gas Density and Molar Mass

$$n = \frac{\text{mass}}{\text{molar mass}}$$

substituting & rearranging

$$P = \frac{\text{mass}}{\text{molar mass}} \times \frac{RT}{V}$$

mass / V = d (the density in g / L)

substituting & rearranging

$$\text{molar mass} = \frac{dRT}{P}$$

Examples

- What is the density of ammonia at 23ºC and 735 torr?

Gases and Stoichiometry

- **Reactions happen in moles**
- **At Standard Temperature and Pressure (STP, 0ºC and 1 atm) 1 mole of gas occuppies 22.42 L.**

$$\frac{14.0\ g\ N_2}{} \Big| \frac{1\ mol\ N_2}{28\ g\ N_2} \Big| \frac{2\ mol\ NH_3}{1\ mol\ N_2} \Big| \frac{22.4\ liters\ NH_3}{1\ mol\ NH_3} = 22.4\ L$$

GIVEN	MVG	MR	FW		ANSWER
$\left(\frac{500\ L\ NH_3\ STP}{1}\right)$	$\left(\frac{1\ mol}{22.4\ L\ STP}\right)$	$\left(\frac{1\ mol\ Mg(OH)_2}{2\ mol\ NH_3}\right)$	$\left(\frac{58.3\ g}{mol}\right)$	= 650.67g	= 651g

Examples

- **Mercury can be achieved by the following reactionWhat volume of oxygen gas can be produced from 4.10 g of mercury (II) oxide at STP?**

Examples

- Consider the following reaction What volume of NO at 1.0 atm and 1000°C can be produced from 10.0 L of NH_3 and excess O_2 at the same temperture and pressure?

$$4NH_3(g) + 5O_2(g) \rightarrow 4\ NO(g) + 6H_2O(g)$$

- What volume of O_2 measured at STP will be consumed when 10.0 kg NH_3 is reacted?

The Same reaction

$$4NH_3(g) + 5\,O_2(g) \rightarrow 4\,NO(g) + 6H_2O(g)$$

- **What mass of H_2O will be produced from 65.0 L of O_2 and 75.0 L of NH_3 both measured at STP?**
- **What volume Of NO would be produced?**
- **What mass of NO is produced from 500. L of NH3 at 250.0ºC and 3.00 atm?**

Dalton's Law

- The total pressure in a container is the sum of the pressure each gas would exert if it were alone in the container.
- The total pressure is the sum of the partial pressures.
- $P_{Total} = P_1 + P_2 + P_3 + P_4 + P_5$ ...
- For each P = nRT/V

Dalton's Law

- $P_{Total} = \frac{n_1RT}{V} + \frac{n_2RT}{V} + \frac{n_3RT}{V} + \ldots$
- In the same container R, T and V are the same.
- $P_{Total} = (n_1 + n_2 + n_3 + \ldots)\frac{RT}{V}$
- $\frac{Pp}{Pt} = \frac{Np}{Nt}$

Examples

- **The partial pressure of nitrogen in air is 592 torr. Air pressure is 752 torr, what is the mole fraction of nitrogen?**

Vapor Pressure

- **Water evaporates!**
- **When that water evaporates, the vapor has a pressure.**
- **Gases are often collected over water so the vapor. pressure of water must be subtracted from the total pressure.**

P = Ptotal – Pvapor water

Kinetic Molecular Theory

- Theory tells why the things happen.
- Explains why ideal gases behave the way they do.

1. All particles are in constant, random motion
1. Temperature is a measure of KE
1. Pressure is due to collisions of gas particles with the walls of the container.

Kinetic Molecular Theory

- Increased Temperature causes more collisions as well as harder collisions.
- Some particles are moving fast, some are moving slowly (depends on mass)

- Ideal gas is a gas in which the particles of the gas do not interact with each other, but only with the walls of the container in which the gas is captured, and only in 100% elastic collisions.

  In real gas the molecules or atoms do interact with each other, rejecting or attracting each other (Van der Waals forces) and transferring energy to each other. Additionally, the collisions with the walls are not always that perfect.

Van der Waals

- Equation has been taken out of AP exam, but you need to know this:
- Predicts the behavior of real gases at low temperature and/or high presure
  - Deviates from ideal behavior under the following conditions
    - Large molecular mass
    - Low volume/high pressure
    - Low temperature

Under what conditions does a gas behave more like a Real gas then an Ideal Gas?

At Low Temperatures and High Pressure.

Velocity

- Average increases as temperature increases.
- Spread increases as temperature increases.

Effusion

- **Passage of gas through a small hole, into a vacuum.**
- **The effusion rate measures how fast this happens.**
- **Graham's Law the rate of effusion is inversely proportional to the square root of the mass of its particles.**

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Diffusion

- **The spreading of a gas through a room.**
- **Slow considering molecules move at 100’s of meters per second.**