Chapter 14

The Gas Laws

The Gas Laws

- **Describe HOW gases behave.**
- **Can be predicted by the theory.**
  - The Kinetic Theory
- **Amount of change can be calculated with mathematical equations.**

The effect of adding gas.

- **When we blow up a balloon we are adding gas molecules.**
- **Doubling the number of gas particles doubles the pressure**

(of the same volume at the same temperature).

4 things

- **In order to completely describe a gas you need to measure 4 things**

1. Pressure
2. Temperature
3. Volume
4. Number of particles

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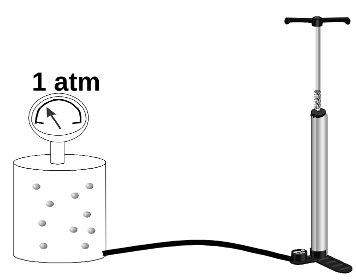
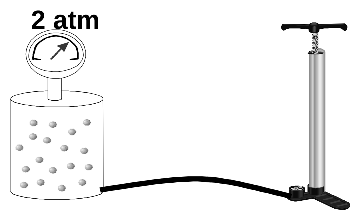
Pressure and the number of molecules are directly related

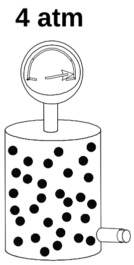
- **More molecules means more collisions**
- **Fewer molecules means fewer collisions.**

- **If you double the number of molecules**

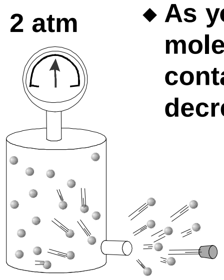


- If you double the number of molecules
- You double the pressure.

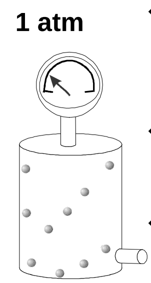




- As you remove molecules from a container



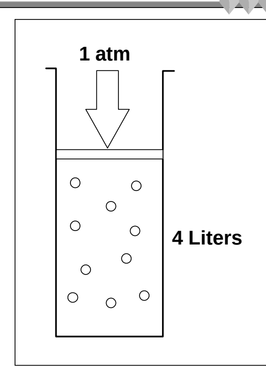
- As you remove molecules from a container the pressure decreases



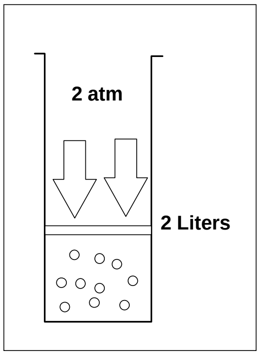
- As you remove molecules from a container the pressure decreases
- Until the pressure inside equals the pressure outside
- Molecules naturally move from high to low pressure

Changing the size of the container

- In a smaller container molecules have less room to move
- Hit the sides of the container more often
- As volume decreases pressure increases.



- As the pressure on a gas increases

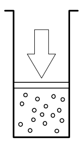


- As the pressure on a gas increases the volume decreases
- Pressure and volume are inversely related

Temperature

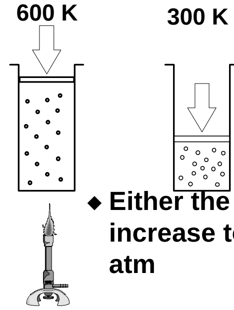
- Raising the temperature of a gas increases the pressure if the volume is held constant.
- The molecules hit the walls harder.
- The only way to increase the temperature at constant pressure is to increase the volume.

300 K



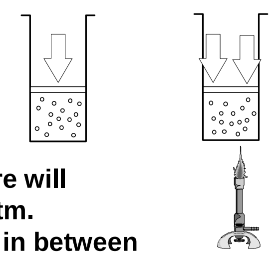
- If you start with 1 liter of gas at 1 atm pressure and 300 K
- and heat it to 600 K one of 2 things happens

600 K 300 K



- Either the volume will increase to 2 liters at 1 atm

300 K 600 K



•Or the pressure will increase to 2 atm.
•Or someplace in between

Ideal Gases

- In this chapter we are going to assume the gases behave ideally
- Does not really exist
  - makes the math easier
  - close approximation.
- Assume particles have no volume
- Assume no attractive forces between molecules

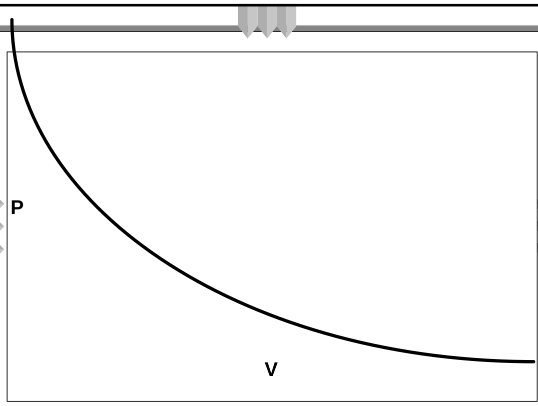
Ideal Gases

- There are no gases for which this is true.
- Real gases behave this way at high temperature and low pressure.

Boyle's Law

- At a constant temperature pressure and volume are inversely related
- As one goes up the other goes down
- $P \times V = \mathcal{K}$ ($\mathcal{K}$ is some constant)
- Easier to use $P_1 \times V_1 = P_2 \times V_2$

Graph



Example

- A balloon is filled with 25 L of air at 1.0 atm pressure. If the pressure is changed to 1.5 atm what is the new volume?

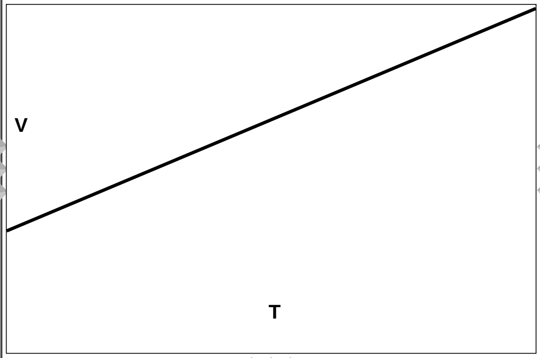
Example

- A balloon is filled with 73 L of air at 1.3 atm pressure. What pressure is needed to change to volume to 43 L?

Charles' Law

- The volume of a gas is directly proportional to the Kelvin temperature if the pressure is held constant.
- $V = \mathcal{K} \times T$ ($\mathcal{K}$ is some constant)
- $\frac{V}{T} = \mathcal{K}$
- $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Graph



Example

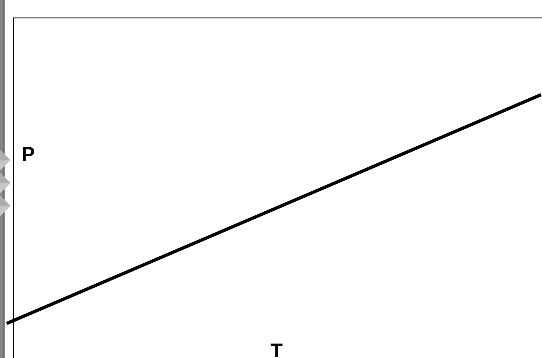
- What is the temperature of a gas that is expanded from 2.5 L at 25ºC to 4.1L at constant pressure.

Example

- What is the final volume of a gas that starts at 8.3 L and 17ºC and is heated to 96ºC?

Gay Lussac's Law

- The temperature and the pressure of a gas are directly related at constant volume.
- $P = \mathcal{K} \times T$ ($\mathcal{K}$ is some constant)
- $\frac{P}{T} = \mathcal{K}$
- $\frac{P_1}{T_1} = \frac{P_2}{T_2}$



Examples

- What is the pressure inside a 0.250 L can of deodorant that starts at 25ºC and 1.2 atm if the temperature is raised to 100ºC?
- At what temperature will the can above have a pressure of 2.2 atm?

Animation

Putting the pieces together

- The Combined Gas Law Deals with the situation where only the number of molecules stays constant.
- $\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$
- Lets us figure out one thing when two of the others change.

- The combined gas law contains all the other gas laws!
- If the temperature remains constant.

$$P_1 \times V_1 = P_2 \times V_2$$

Boyle's Law

- The combined gas law contains all the other gas laws!
- If the pressure remains constant.

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

Charles' Law

- The combined gas law contains all the other gas laws!
- If the volume remains constant.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Gay-Lussac's Law

Examples

- A 15 L cylinder of gas at 4.8 atm pressure at 25ºC is heated to 75ºC and compressed to 17 atm. What is the new volume?

Examples

- If 6.2 L of gas at 723 mm Hg at 21ºC is compressed to 2.2 L at 4117 mm Hg, what is the temperature of the gas?

The Fourth Part

- Avagadro's Hypothesis
- V is proportional to number of molecules at constant T and P.
- V is proportional to moles.
- V = $\mathcal{K}$n (n is the number of moles.
- Gets put into the combined gas law
- $\frac{P_1 \times V_1}{n_1 \times T_1} = \frac{P_2 \times V_2}{n_1 \times T_2} = R$

The Ideal Gas Law

- P x V = n x R x T
- Pressure times Volume equals the number of moles times the Ideal Gas Constant (R) times the temperature in Kelvin.
- This time R does not depend on anything, it is really constant

The Ideal Gas Constant

- $R = 0.0821 \frac{(L\ atm)}{(mol\ K)}$
- $R = 62.4 \frac{(L\ mm\ Hg)}{(K\ mol)}$
- $R = 8.31 \frac{(L\ kPa)}{(K\ mol)}$

The Ideal Gas Law

- PV = nRT
- We now have a new way to count moles of a gas. By measuring T, P, and V.
- We aren't restricted to STP.
- n = PV/RT
- Nothing is required to change,
  - No 1's and 2's

Example

- How many moles of air are there in a 2.0 L bottle at 19ºC and 747 mm Hg?

Example

- What is the pressure exerted by 1.8 g of H_2 gas exert in a 4.3 L balloon at 27ºC?

Density

- The Molar mass of a gas can be determined by the density of the gas.
- $D = \frac{mass}{Volume} = \frac{m}{V}$
- $\text{Molar mass} = \frac{mass}{Moles} = \frac{m}{n}$
- $n = \frac{PV}{RT}$

- $\text{Molar Mass} = \frac{m}{(PV/RT)}$
- $\text{Molar mass} = \frac{m\,RT}{V\;\;P}$
- $\text{Molar mass} = \frac{DRT}{P}$

Stoichiometry and Gases

At STP

- At STP determining the amount of gas required or produced is easy.
- 22.4 L = 1 mole

Not At STP

- Chemical reactions happen in MOLES.
- If you know how much gas - change it to moles
- Use the Ideal Gas Law $n = PV/RT$
- If you want to find how much gas - use moles to figure out volume $V = nRT/P$
- Use the equation in place of 22.4 L

Example

- HCl(g) can be formed by the following reaction
- $2NaCl(aq) + H_2SO_4(aq) \rightarrow 2HCl(g) + Na_2SO_4(aq)$
- What mass of NaCl is needed to produce 340 mL of HCl at 1.51 atm at 20ºC?

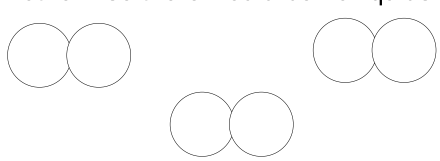
Example

- $2NaCl(aq) + H_2SO_4\ (aq) \rightarrow 2HCl(g) + Na_2SO_4\ (aq)$
- **What volume of HCl gas at 25ºC and 715 mm Hg will be generated if 10.2 g of NaCl react?**

Real Gases

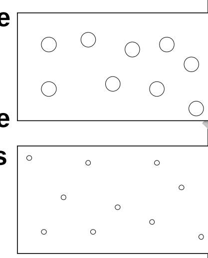
Ideal Gases don't exist

- **Molecules do take up space**
  - All matter has volume
- **There are attractive forces**
  - otherwise there would be no liquids



Real Gases behave like Ideal Gases

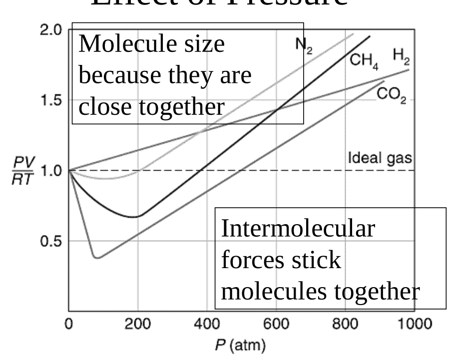
- **When the molecules are far apart**
- **They take a smaller percentage of the space**
- **Ignoring their volume is reasonable**
- **This is at low pressure**



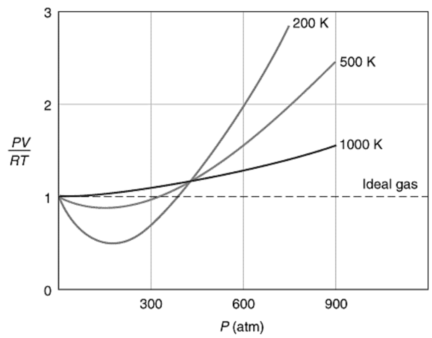
Real Gases behave like Ideal gases when

- **When molecules are moving fast.**
- **Molecules are not next to each other very long**
- **Attractive forces can't play a role.**
- **At high temp.**
- **Far above boiling point.**

Effect of Pressure



Effect of Temperature



Van der Waal's equation

Corrected Pressure

Corrected Volume

- a is a number that depends on how much the molecules stick to each other
- b is a number that determined by how big the molecules are

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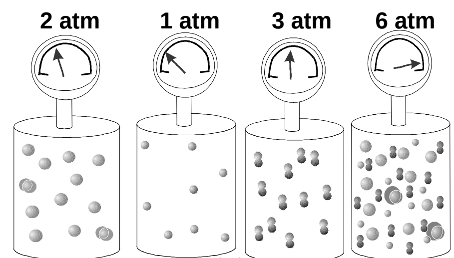
Dalton's Law of Partial Pressures

- The total pressure inside a container is equal to the sum of the partial pressure due to each gas.
- The partial pressure of a gas is the contribution by that gas hitting the wall.
- $P_{Total} = P_1 + P_2 + P_3 + \ldots$
- For example

- We can find out the pressure in the fourth container
- By adding up the pressure in the first 3



Dalton's Law

- This means that we can treat gases in the same container as if they don't affect each other.
- Figure out their pressures separately
- Add them to get total

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Examples

- What is the total pressure in a balloon filled with air if the pressure of the oxygen is 170 mm Hg and the pressure of nitrogen is 620 mm Hg?

Examples

In a second balloon the total pressure is 1.3 atm. What is the pressure of oxygen if the pressure of nitrogen is 720 mm Hg?

- **There are two containers connected by a valve. One holds 4.0 L of N_2 at 2.0 atm the other holds 2.0 L of O_2 at 8.0 atm.**
- **The valve is opened. What is**
  - The pressure of N_2 ?
  - The pressure of O_2 ?
  - The total pressure?

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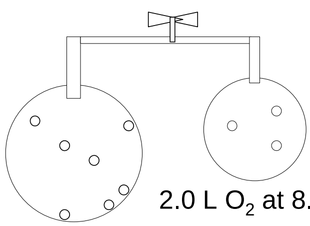

2.0 L O_2 at 8.0 atm

4.0 L N_2 at 2.0 atm

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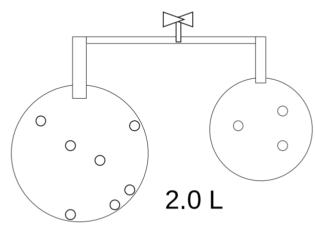

2.0 L

4.0 L

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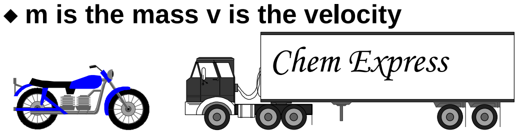
Diffusion

- **Molecules moving from areas of high concentration to low concentration.**
- **Perfume molecules spreading across the room.**
- **Effusion - Gas escaping through a tiny hole in a container.**
- **From high to low concentration**
- **Both depend on the speed of the molecules**

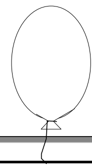
Graham's Law

- **The rate of effusion and diffusion is inversely proportional to the square root of the molar mass of the molecules.**
- **Kinetic energy = 1/2 mv^2**
- **m is the mass v is the velocity**



Graham's Law

- **bigger molecules move slower at the same temp. (by Square root)**
- **Bigger molecules effuse and diffuse slower**
- **Helium effuses and diffuses faster than air - escapes from balloon.**



Grahams Law

$$\frac{\text{effusion rate 1}}{\text{effusion rate 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

$$\frac{\text{Velocity 1}}{\text{Velocity 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

$$\frac{\text{Time 2}}{\text{Time 1}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

Example

- **In a test He effused at 3.5 moles/minute. How fast would N_2 effuse in the same conditions?**

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Example

- **H_2 effused 2.82 times faster than an unknown gas. What was the molar mass of the unknown gas.**

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Example

- **It took an unknown gas 89.3 seconds to effuse through a hole. Oxygen effused through the same hole in 60 seconds. What is the molar mass of the gas?**

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