

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

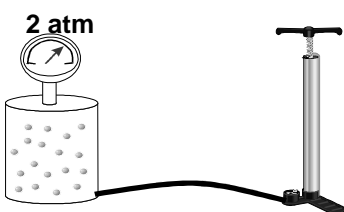
Pressure and the number of molecules are directly related

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

- ◆ If you double the number of molecules

1 atm

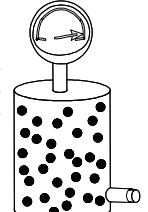
- ◆ If you double the number of molecules
- ◆ You double the pressure.



2 atm

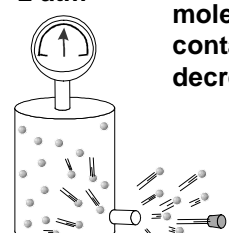
4 atm

- ◆ As you remove molecules from a container



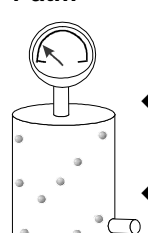
2 atm

- ◆ As you remove molecules from a container the pressure decreases



1 atm

- ◆ 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



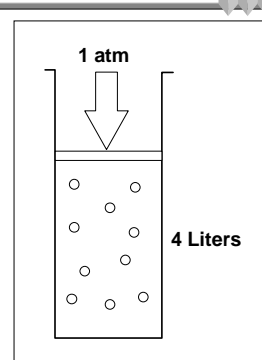
Changing the size of the container

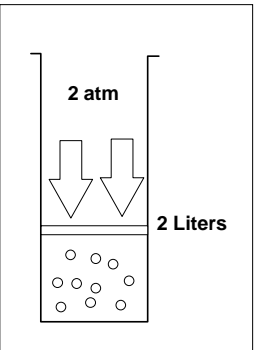
- ◆ In a smaller container molecules have less room to move
- ◆ Hit the sides of the container more often
- ◆ As volume decreases pressure increases.

1 atm

4 Liters

- ◆ As the pressure on a gas increases



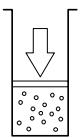


- ◆ As the pressure on a gas increases the volume decreases
- ◆ Pressure and volume are inversely related

Temperature

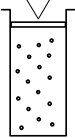
- ◆ Raising the temperature of a gas increases the pressure if the volume is held constant.
- ◆ The molecules hit the walls harder.
- ◆ The only way to increase the temperature at constant pressure is to increase the volume.

300 K

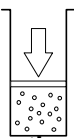


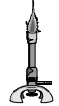
- ◆ If you start with 1 liter of gas at 1 atm pressure and 300 K
- ◆ and heat it to 600 K one of 2 things happens

600 K



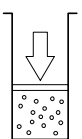
300 K



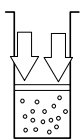



- ◆ Either the volume will increase to 2 liters at 1 atm

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

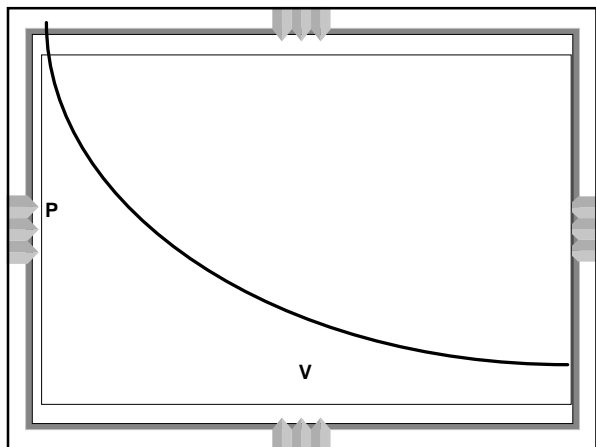
Ideal Gases

- ◆ 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
- ◆ $P \times V = \kappa$ (κ 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?

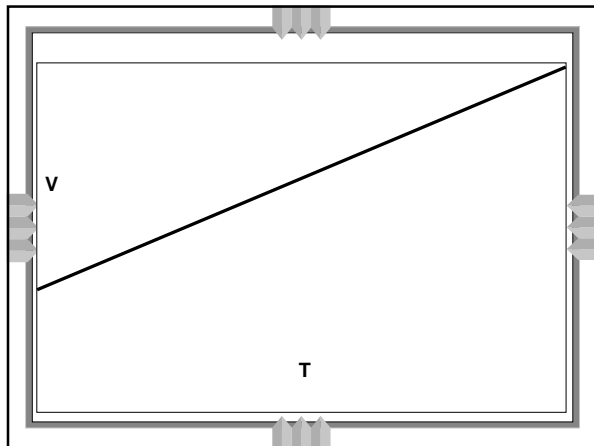
Example

- ◆ 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.
- ◆ $V = \kappa \times T$ (κ is some constant)
- ◆ $\frac{V}{T} = \kappa$
- ◆ $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Graph



Example

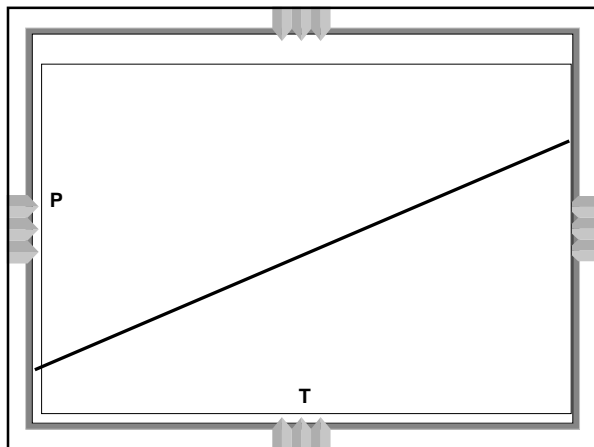
- ◆ What is the temperature of a gas that is expanded from 2.5 L at 25°C to 4.1L at constant pressure.

Example

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

Gay Lussac's Law

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



Examples

- ◆ What is the pressure inside a 0.250 L can of deodorant that starts at 25°C and 1.2 atm if the temperature is raised to 100°C?
- ◆ 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.

$$\frac{P_1 \times V_1}{\boxed{}} = \frac{P_2 \times V_2}{\boxed{}}$$

Boyle's Law

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

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

Charles' Law

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

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

Gay-Lussac's Law

Examples

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

Examples

- ◆ If 6.2 L of gas at 723 mm Hg at 21°C 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.
- ◆ $V = \kappa n$ (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

- ◆ $P \times V = n \times R \times T$
- ◆ 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

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

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 19°C and 747 mm Hg?

Example

- ◆ What is the pressure exerted by 1.8 g of H₂ gas exert in a 4.3 L balloon at 27°C?

Density

- ◆ The Molar mass of a gas can be determined by the density of the gas.

$$\text{◆ } D = \frac{\text{mass}}{\text{Volume}} = \frac{m}{V}$$

$$\text{◆ Molar mass} = \frac{\text{mass}}{\text{Moles}} = \frac{m}{n}$$

$$\text{◆ } n = \frac{PV}{RT}$$

$$\text{◆ Molar Mass} = \frac{m}{(PV/RT)}$$

$$\text{◆ Molar mass} = \frac{m RT}{V P}$$

$$\text{◆ Molar mass} = \frac{DRT}{P}$$

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
- ◆ $2\text{NaCl(aq)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow 2\text{HCl(g)} + \text{Na}_2\text{SO}_4(\text{aq})$
- ◆ What mass of NaCl is needed to produce 340 mL of HCl at 1.51 atm at 20°C?

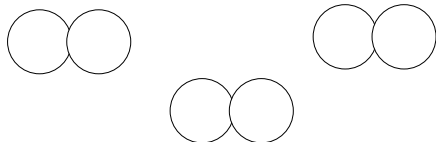
Example

- ◆ $2\text{NaCl}(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow 2\text{HCl}(\text{g}) + \text{Na}_2\text{SO}_4(\text{aq})$
- ◆ What volume of HCl gas at 25°C and 715 mm Hg will be generated if 10.2 g of NaCl react?

Real Gases

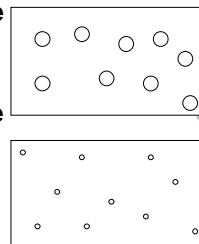
Ideal Gases don't exist

- ◆ **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

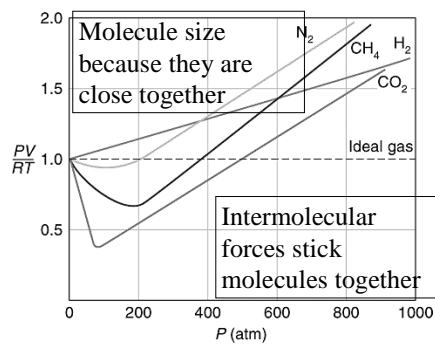
- ◆ **When the molecules are far apart**
- ◆ **They take a smaller percentage of the space**
- ◆ **Ignoring their volume is reasonable**
- ◆ **This is at low pressure**

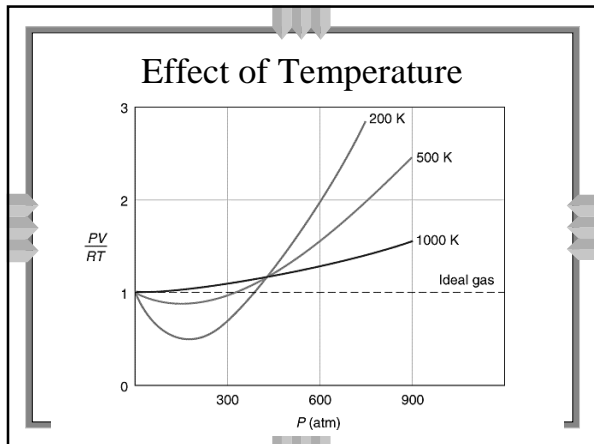


Real Gases behave like Ideal gases when

- ◆ **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





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

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.
- ◆ $P_{\text{Total}} = P_1 + P_2 + P_3 + \dots$
- ◆ 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

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?

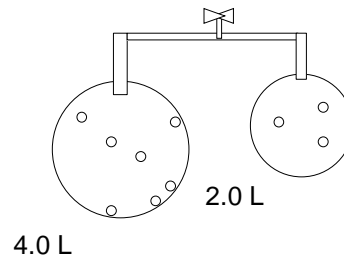
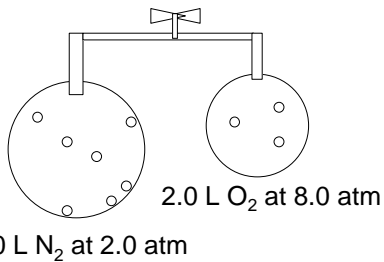
Examples

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?



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

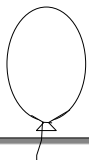
Graham's Law

- ◆ The rate of effusion and diffusion is inversely proportional to the square root of the molar mass of the molecules.
- ◆ Kinetic energy = $\frac{1}{2} mv^2$
- ◆ 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 N₂ effuse in the same conditions?

Example

- ◆ H₂ effused 2.82 times faster than an unknown gas. What was the molar mass of the unknown gas.

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?