

# Gases

## Ideal Gas Law

## Ideal Gas Law

- **Ideal gas:** a gas in which all collisions between atoms or molecules...
  - are *perfectly elastic* (no energy lost)
  - there are *no intermolecular attractive forces*
- Think of an ideal gas as a collection of perfectly **hard spheres** which collide, but which otherwise do not interact with each other. Like billiards!



## Ideal Gas Law

- An ideal gas can be characterized by three variables:
  - pressure (P), with units of **atm**
  - volume (V), with units of **L**
  - temperature (T), with units of **K**
- The relationship between them may be deduced from kinetic molecular theory.



## Kinetic Molecular Theory



- The gas consists of very small particles, all with non-zero mass.
- These molecules are in constant, random motion. The rapidly moving particles constantly collide with the walls of the container.
- The collisions of gas particles with the walls of the container holding them are perfectly elastic.
- The interactions among molecules are negligible. They exert no forces on one another except during collisions.

## Exceptions to the Ideal Gas Law

- at **low temperatures** (close to 0 K)
  - gas molecules have less kinetic energy (move around less) so they *do* attract each other
- at **high pressures** (like many, many atm)
  - gas molecules are forced closer together so that the volume of the gas molecules becomes significant compared to the volume the gas occupies



## Measuring Temperature

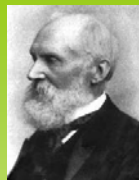
- **Fahrenheit (°F):**
  - based on the freezing and melting point of water
- **Celsius (°C):**
  - based on the freezing and melting point of water, but in base 10
- **Kelvin (K):**
  - based on the idea that at 0 Kelvin or absolute zero, atoms are **NOT** in motion

**Absolute Zero**  
(it's the coolest)

## Converting between Celsius and Kelvin



and Kelvin



$$K = ^\circ C + 273$$

### Example:

The melting point of mercury is  $-39^\circ\text{C}$ . What is this temperature in Kelvin?

## Symbolic Representation Ideal Gas Law

$$PV = nRT$$

- $n$  = number of **moles** of gas
- $R$  = universal gas constant =  $0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$

## Ideal Gas Law

*A few things to know...*

- STP: standard temp and pressure
  - $T = 273.15 \text{ K}$
  - $P = 1.000 \text{ atm}$
- Diatomic molecules:

H O F Br I N Cl

## Ideal Gas Law Example #1

What volume is needed to store 0.050 moles of helium gas at 2.00 atm and 400 K?

## Ideal Gas Law Example #2

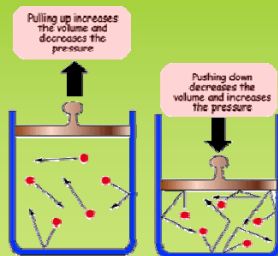
What pressure will be exerted by 20.16 g of hydrogen gas in a 7.5 L cylinder at  $20^\circ\text{C}$ ?

# Gases

Boyle's Law

## Boyle's Law

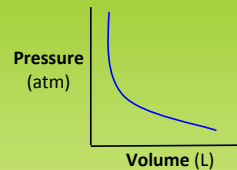
- As the volume of a gas increases, the pressure decreases.
  - Temperature remains constant



In the smaller space, the particles suffer more collisions with the walls of the container - it is this that we measure as pressure exerted by the gas.

## Symbolic Representation of Boyle's Law

As volume increases, pressure decreases



Pressure and temperature are *inversely* proportional

How is this represented in the form of an equation?  $P_1V_1 = P_2V_2$

## Units of Measuring Pressure

- Many ways to measure pressure of a gas.
  - Atmospheres (atm): force per unit area exerted against a surface by the weight of air above that surface in the Earth's atmosphere.
  - mm Hg (mmHg): pressure of 1 mm of mercury in a column of mercury
  - torr equal to 1 mmHg
  - Pascals (Pa) or kilopascals (kPa)



## Boyle's Law Example #1

A balloon contains 20.0 L of helium gas at 1.02 atm. The balloon rises to an altitude where the pressure is only 0.15 atm. What is the volume of the balloon at the higher altitude?

**Step #1:** Qualitative Assessment of the Problem

**Step #2:** Quantitative Assessment of the Problem

## Boyle's Law Example #2

A gas with a volume of 4.00 L at a pressure of 2.02 atm is allowed to expand to a volume of 12.0 L. What is the pressure in the container if the temperature remains constant?

**Step #1:** Qualitative Assessment of the Problem

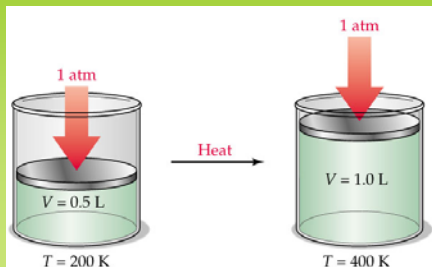
**Step #2:** Quantitative Assessment of the Problem

# Gases

Charles' Law

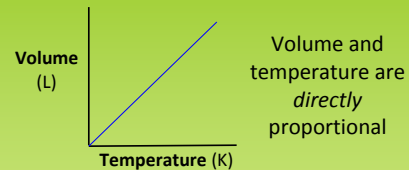
## Charles' Law

- As the temperature of a gas increases, the volume increases.
  - Pressure remains constant



## Symbolic Representation of Charles' Law

As temperature increases, volume increases



Volume and temperature are *directly* proportional

How is this represented in the form of an equation?

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

## Charles' Law Example #1

A balloon has a volume of 1.00 L. The temperature of the room where the balloon is located is 273 K. If the room is heated to 300 K, what happens to the volume of the balloon?

**Step #1:** Qualitative Assessment of the Problem

**Step #2:** Quantitative Assessment of the Problem

## Charles' Law Example #2

Exactly 5.00 L of air at  $-50.0^\circ\text{C}$  is warmed to  $100.0^\circ\text{C}$ . What is the new volume of the air if the pressure remains constant?

**Step #1:** Qualitative Assessment of the Problem

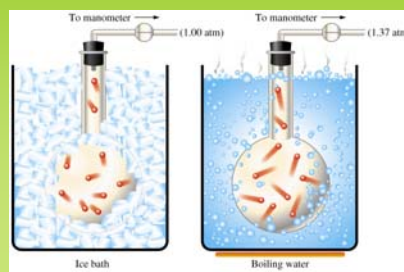
**Step #2:** Quantitative Assessment of the Problem

# Gases

Gay Lussac's Law

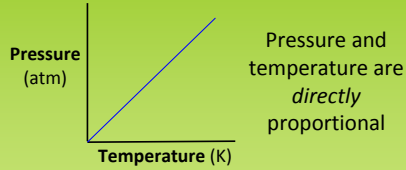
## Gay-Lussac's Law

- As the temperature of a gas increases, the pressure increases.
  - Volume remains constant



## Symbolic Representation of Gay-Lussac's Law

As temperature increases, pressure increases



How is this represented in the form of an equation?

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

## Gay-Lussac's Law Example #1

A sealed cylinder contains nitrogen gas at 9.8 atm and a temperature of 293 K. When the cylinder is left in the sun, the temperature of the gas increases to 323 K. What is the pressure of the gas in the cylinder?

**Step #1:** Qualitative Assessment of the Problem

**Step #2:** Quantitative Assessment of the Problem

## Gay-Lussac's Law Example #2

The pressure in a car tire is 1.95 atm at 27°C. After a long drive, the pressure is 2.22 atm. What is the temperature of the air in the tire? Assume that the volume is constant.

**Step #1:** Qualitative Assessment of the Problem

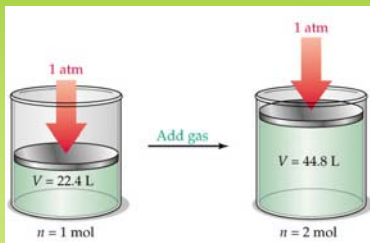
**Step #2:** Quantitative Assessment of the Problem

# Gases

## Avogadro's Law

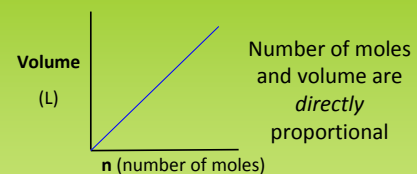
## Avogadro's Law

- As the number of moles of a gas increases, the volume increases.
  - Pressure and temperature remain constant



## Symbolic Representation of Avogadro's Law

As number of moles of gas increases, volume increases



How is this represented in the form of an equation?

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

# Gases

## Dalton's Law of Partial Pressures

## Partial Pressures

What do you think the total pressure of the mixture below is?

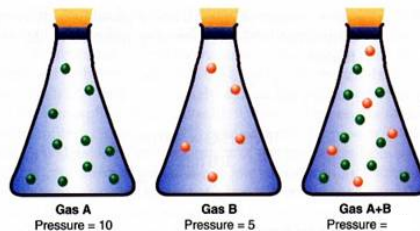


Figure 3-1. Dalton's law.

## Dalton's Law of Partial Pressures

"The total pressure of a mixture of gases equals the sum of the pressure that each would exert if it were present alone." (p. 410, your textbook)

Each gas in a mixture behaves independently. So...

$$P_{\text{total}} = P_1 + P_2 + P_3 + \dots$$

## Mole Fractions & Partial Pressures

To determine the partial pressure of a gas...

$$P_1 = \left( \frac{n_1}{n_{\text{total}}} \right) P_{\text{total}}$$

## Mole Fraction Example

A study of the effects of certain gases on plant growth requires a synthetic atmosphere composed of 1.50 mol  $\text{CO}_2$ , 18.0 mol  $\text{O}_2$ , and 80.5 mol Ar. Calculate the partial pressure of  $\text{O}_2$  if the total pressure of this atmosphere is 745 torr.

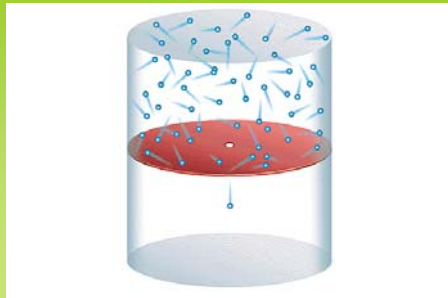
$$P_1 = \left( \frac{n_1}{n_{\text{total}}} \right) P_{\text{total}}$$

# Gases

## Graham's Law of Effusion

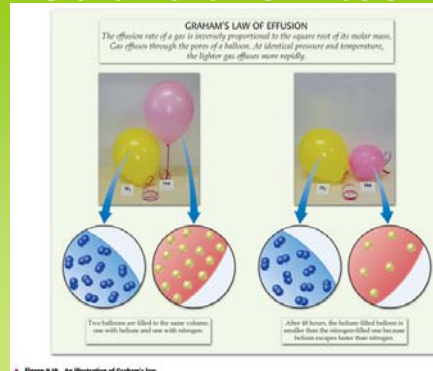
## Effusion

**Effusion:** the escape of gas through a small hole



[http://www.chem.ufl.edu/~rl/2041\\_u00/lectures/lec\\_d.html](http://www.chem.ufl.edu/~rl/2041_u00/lectures/lec_d.html)

## Graham's Law of Effusion



## Graham's Law of Effusion

$$\frac{\text{Rate}_1}{\text{Rate}_2} = \sqrt{\frac{\text{molar mass}_2}{\text{molar mass}_1}}$$

## Graham's Law of Effusion Example

If equal amounts of helium and argon are placed in a porous container and allowed to effuse, which gas will escape faster and how much faster?

$$\frac{\text{Rate}_1}{\text{Rate}_2} = \sqrt{\frac{\text{molar mass}_2}{\text{molar mass}_1}}$$

## Graham's Law of Effusion Example

It takes 354 seconds for a sample of xenon to effuse through a small hole. Under the same conditions, how long will it take nitrogen to effuse?

$$\frac{\text{Rate}_1}{\text{Rate}_2} = \sqrt{\frac{\text{molar mass}_2}{\text{molar mass}_1}}$$

# Gases

## Molar Volume

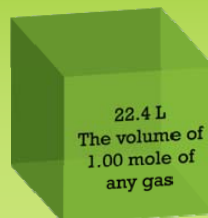
## Molar Volume

- At standard temp and pressure (STP) 1 mole of any gas occupies 22.4 L

1 mole = 22.4 L (use this as a conversion factor)

### Example:

How many liters will 48.6 grams of  $\text{CO}_2$  occupy at STP?



# Gases

## Gas Stoichiometry

## Gas Stoichiometry

- Works just like stoichiometry we have used before
  - Write and balance the equation first!
  - Start with your known value and use dimensional analysis to find your unknown quantity – **don't forget the molar ratios!**
  - Remember: 1 mole of any gas = 22.4 L

## Gas Stoich Example #1

An excess of hydrogen gas reacts with 14.0 g of nitrogen gas to produce ammonia ( $\text{NH}_3$ ). How many liters of ammonia are produced at STP?

## Gas Stoich Example #2

How many liters of  $\text{O}_{2(g)}$  at STP are evolved when 3.25 g of  $\text{KNO}_{3(s)}$  decomposes to  $\text{KNO}_{2(s)}$  and  $\text{O}_{2(g)}$  ?