Gases

Chapter 12

Properties of Gases

- Expand to completely fill their container
- Take the Shape of their container
- Low Density
  - much less than solid or liquid state
- Compressible
- Mixtures of gases are always homogeneous
- Fluid

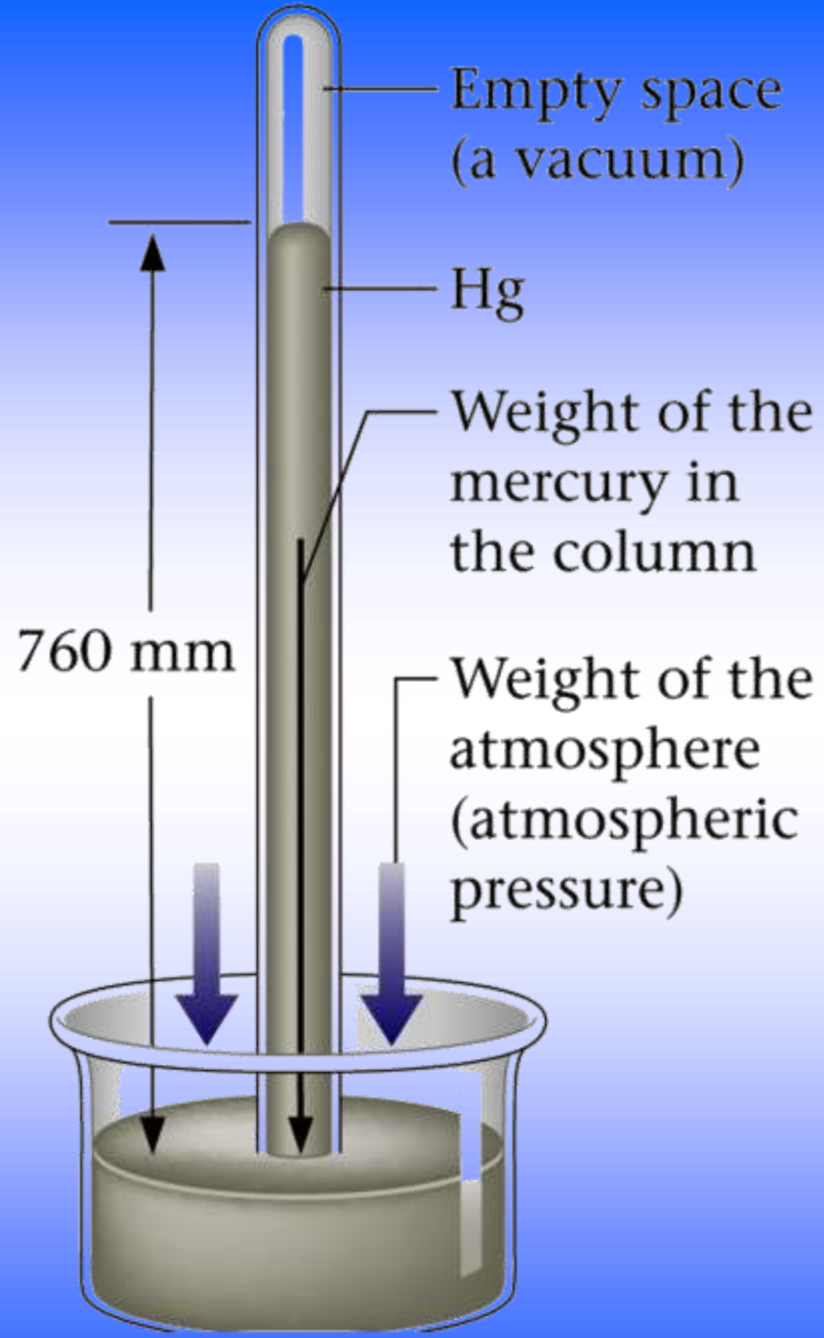
Gas Pressure

- **Pressure** = total force applied to a certain area
  - larger force = larger pressure
  - smaller area = larger pressure
- Gas pressure caused by gas molecules colliding with container or surface
- More forceful collisions or more frequent collisions mean higher gas pressure

Air Pressure

- Constantly present when air present
- Decreases with altitude
  - less air
- Varies with weather conditions
- Measured using a **barometer**
  - Column of mercury supported by air pressure
  - Longer mercury column supported = higher pressure
  - Force of the air on the surface of the mercury balanced by the pull of gravity on the column of mercury

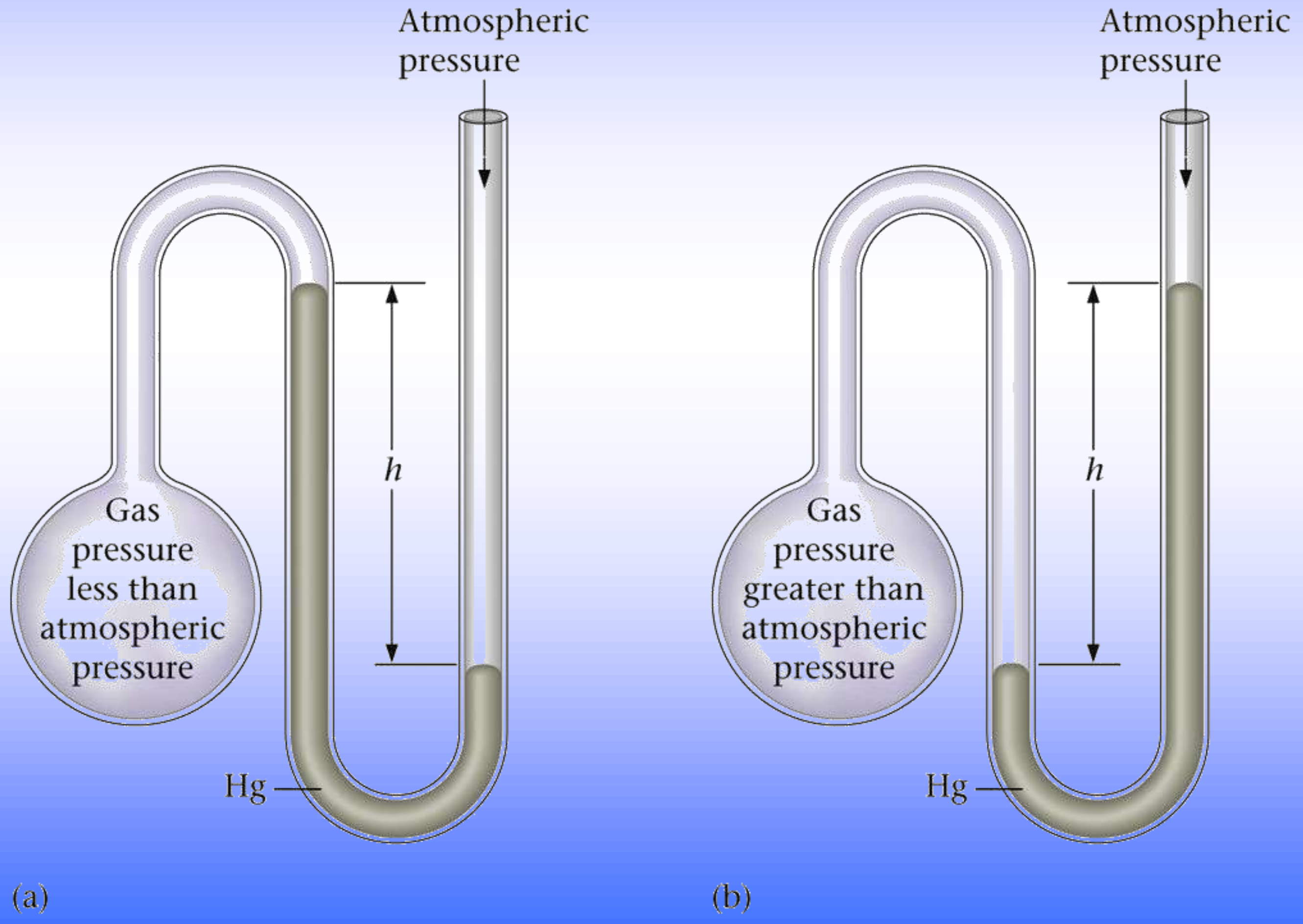
When a glass tube is filled with mercury and inverted in a dish of mercury at sea level, the mercury flows out of the tube until a column approximately 760 mm high remains.



Measuring Pressure of a Trapped Gas

- Use a **manometer**
- Open-end manometer
  - if gas end lower than open end, $P_{gas} = P_{air}$ + diff. in height of Hg
  - if gas end higher than open end, $P_{gas} = P_{air}$ – diff. in height of Hg

A device (called a manometer) for measuring the pressure of a gas in a container.



Units of Gas Pressure

- atmosphere (atm)
- height of a column of mercury (mm Hg, in Hg)
- torr
- Pascal (Pa)
- pounds per square inch (psi, lbs./in^2)
- 1.000 atm = 760.0 mm Hg = 29.92 in Hg = 760.0 torr = 101,325 Pa = 101.325 kPa = 14.69 psi

Boyle's Law

- Pressure is inversely proportional to Volume
  - constant T and amount of gas
  - graph P vs V is curve
  - graph P vs 1/V is straight line
- as P increases, V decreases by the same factor
- $P \times V = \text{constant}$
- $P_1 \times V_1 = P_2 \times V_2$

Table 12.1 A Sample of Boyle's Observations (moles of gas and temperature both constant)

Experiment	Pressure (in Hg)	Volume (in.^{3})	*Pressure × Volume (in Hg) × (in.^{3})* Actual	Rounded*
1	29.1	48.0	1396.8	1.40×10^3
2	35.3	40.0	1412.0	1.41×10^3
3	44.2	32.0	1414.4	1.41×10^3
4	58.2	24.0	1396.8	1.40×10^3
5	70.7	20.0	1414.0	1.41×10^3
6	87.2	16.0	1395.2	1.40×10^3
7	117.5	12.0	1410.0	1.41×10^3

*Three significant figures are allowed in the product because both of the numbers that are multiplied together have three significant figures.

Example

What is the new volume if a 1.5 L sample of freon-12 at 56 torr is compressed to 150 torr?

- Write down the given amounts

  P_1 = 56 torr P_2 = 150 torr

  V_1 = 1.5 L. V_2 = ? L

- Convert values of like quantities to the same units

  both Pressure already in torr

  value of V_2 will come out in L

Example

What is the new volume if a 1.5 L sample of freon-12 at 56 torr is compressed to 150 torr?

- Choose the correct Gas Law

  Since we are looking at the relationship between pressure and volume we use Boyle's Law

  $P_1 \times V_1 = P_2 \times V_2$

- Solve the equation for the unknown variable

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

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

Example

What is the new volume if a 1.5 L sample of freon-12 at 56 torr is compressed to 150 torr?

- Plug in the known values and calculate the unknown

P_1 = 56 torr P_2 = 150 torr

V_1 = 1.5 L. V_2 = ? L

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

$$\frac{56 \text{ torr}}{150 \text{ torr}} \times 1.5 \text{ L} = 0.56 \text{ L}$$

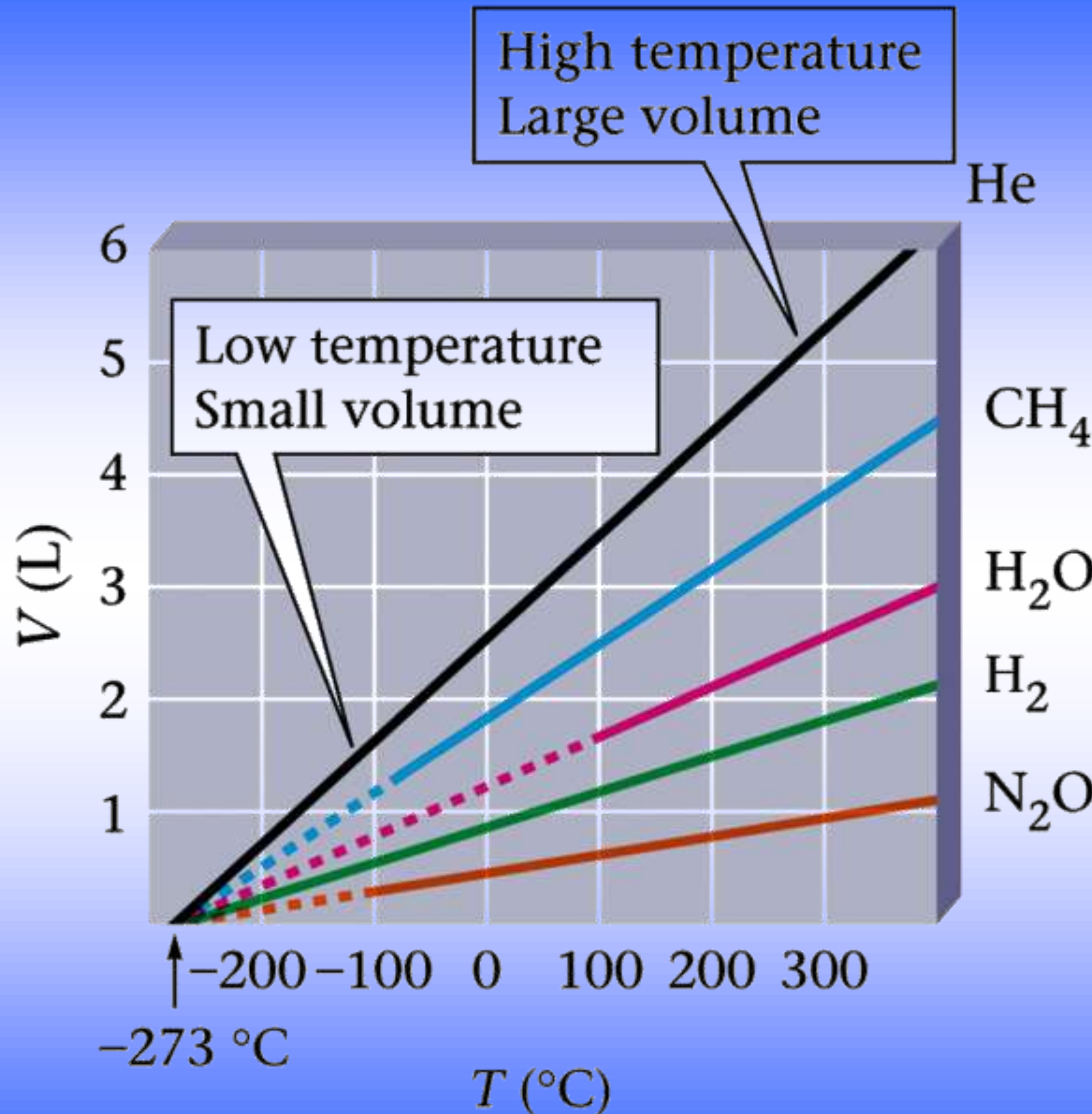
Absolute Zero

- Theoretical ***temperature*** at which a gas would have zero volume and no pressure
  - calculated by extrapolation
- 0 K = -273.15 °C = -459 °F
- Kelvin T = Celsius T + 273.15
- Never attainable
  - though we've gotten real close!
- All gas law problems use Kelvin temperature scale!

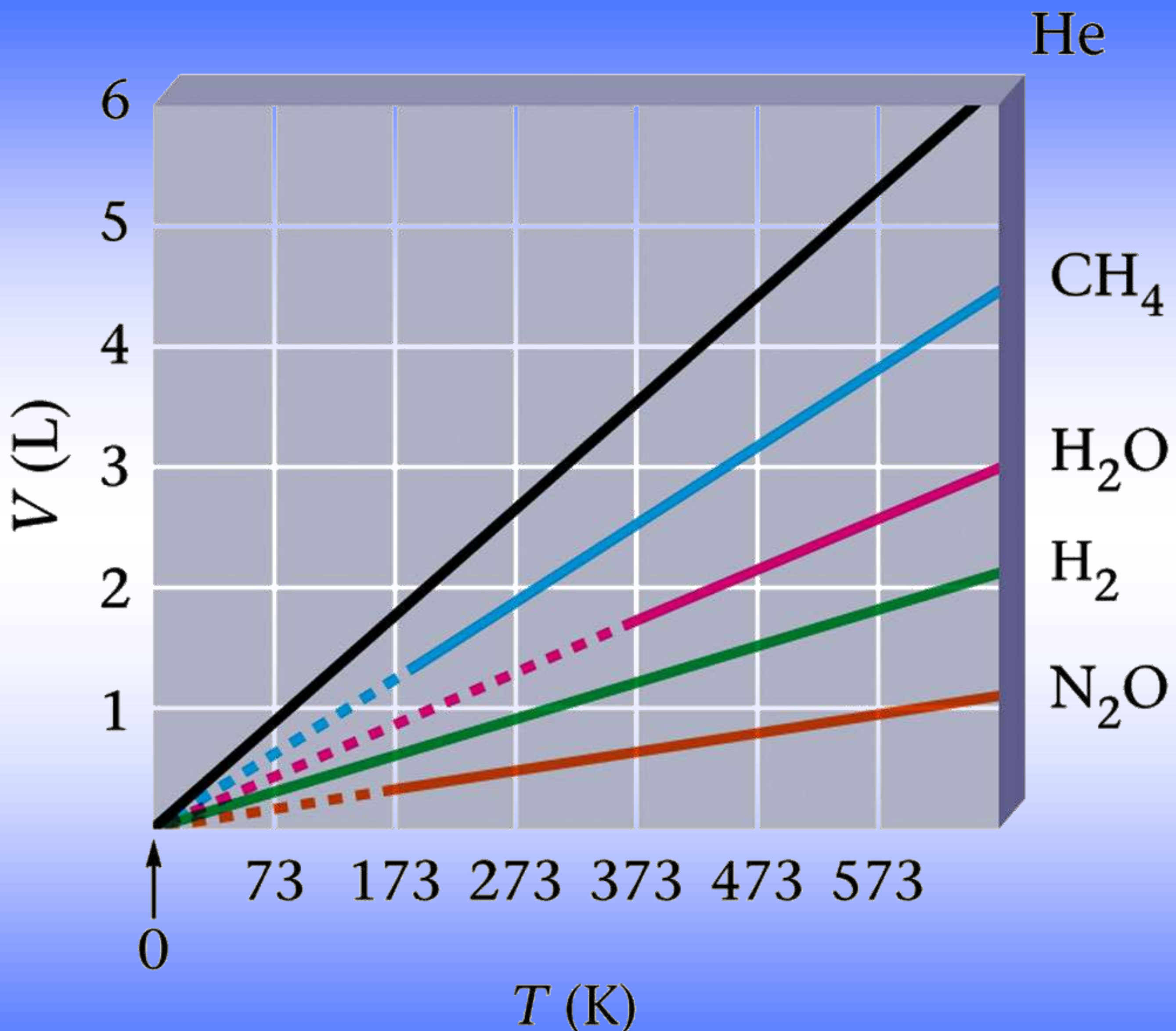
Charles' Law

- Volume is directly proportional to Temperature
  - constant P and amount of gas
  - graph of V vs T is straight line
- as T increases, V also increases
- V = constant x T
  - if T measured in kelvin
- $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Plots of V (L) versus *T* (°C) for several gases.



Plots of *V* versus *T* as in Figure 12.7, except that here the Kelvin scale is used for temperature.



Avogadro's Law

- Volume directly proportional to the number of gas molecules
  - **V** = constant x **n (moles)**
  - Constant P and T
  - More gas molecules = larger volume

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

- Count number of gas molecules by **moles**
- One mole of any ideal gas occupies 22.414 L at standard conditions - **molar volume**
- Equal volumes of gases contain equal numbers of molecules
  - ***It doesn't matter what the gas is!***

Ideal Gas Law

- By combing the proportionality constants from the gas laws we can write a general equation
- **R** is called the **gas constant**
- The value of **R** depends on the units of P and V
  - Generally use R = 0.08206 when P in atm and V in L
- Use the ideal gas law when have gas at one condition
- Most gases obey this law when pressure is low (at or below 1 atm) and temperature is high (above 0°C)
- If a gas changes some conditions, the unchanging conditions drop out of the equation

$$PV = nRT$$

Combined Gas Law

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

Dalton's Law

- The total pressure of a mixture of gases equals the sum of the pressures each gas would exert independently
  - **Partial pressures** is the pressure a gas in a mixture would exert if it were alone in the container
  - $\mathbf{P_{total} = P_{gas\ A} + P_{gas\ B} + \ldots}$
- Particularly useful for determining the pressure a dry gas would have after it is collected over water
  - $\mathbf{P_{air} = P_{wet\ gas} = P_{dry\ gas} + P_{water\ vapor}}$
  - $P_{water\ vapor}$ depends on the temperature, look up in table

Partial Pressures

The partial pressure of each gas in a mixture can be calculated using the Ideal Gas Law

for gases A and B in a mixture

$$P_A = \frac{n_A \times R \times T}{V} \qquad P_B = \frac{n_B \times R \times T}{V}$$

the temperature and volume of everything in the mixture are the same

$$n_{total} = n_A + n_B$$

$$P_{total} = P_A + P_B = \frac{n_{total} \times R \times T}{V}$$

Kinetic - Molecular Theory

- The properties of solids, liquids and gases can be explained based on the speed of the molecules and the attractive forces between molecules
- In **solids,** the molecules have no translational freedom, they are held in place by strong attractive forces
  - May only vibrate

Kinetic - Molecular Theory

- In **liquids,** the molecules have some translational freedom, but not enough to escape their attraction for neighboring molecules
  - They can slide past one another, rotate as well as vibrate
- In **gases,** the molecules have “complete” freedom from each other, they have enough energy to overcome “all” attractive forces
- Kinetic energy depends only on the temperature

Describing a Gas

- Gases are composed of tiny particles
- The particles are small compared to the average space between them
  - Assume the molecules do not have volume
- Molecules constantly and rapidly moving in a straight line until they bump into each other or the wall
  - Average kinetic energy proportional to the temperature
  - Results in gas pressure
- Assumed that the gas molecules attraction for each other is negligible

Postulates of the Kinetic Molecular Theory of Gases

1. Gases consist of tiny particles (atoms or molecules).
2. These particles are so small, compared with the distances between them, that the volume (size) of the individual particles can be assumed to be negligible (zero).
3. The particles are in constant random motion, colliding with the walls of the container. These collisions with the walls cause the pressure exerted by the gas.
4. The particles are assumed not to attract or to repel each other.
5. The average kinetic energy of the gas particles is directly proportional to the Kelvin temperature of the gas.

Gas Properties Explained

- Gases have indefinite shape and volume because the freedom of the molecules allows them to move and fill the container they're in
- Gases are compressible and have low density because of the large spaces between the molecules

The Meaning of Temperature

- Temperature is a measure of the <u>average</u> kinetic energy of the molecules in a sample
  - Not all molecules have same kinetic energy
- Kinetic energy is directly proportional to the Kelvin Temperature
  - average speed of molecules increases as the temperature increase

Pressure and Temperature

- As the temperature of a gas increases, the average speed of the molecules increases
- the molecules hit the sides of the container with more force (on average)
- the molecules hit the sides of the container more frequently
- the net result is an ***increase in pressure***

Volume and Temperature

- In a rigid container, raising the temperature increases the pressure
- For a cylinder with a piston, the pressure outside and inside stay the same
- To keep the pressure from rising, the piston moves out increasing the volume of the cylinder
  - as volume increases, pressure decreases

Gas Stoichiometry

- Use the general algorithms discussed previously to convert masses or solution amounts to moles
- Use gas laws to convert amounts of gas to moles
  - or visa versa