Introductory Chemistry, 2nd Edition
Nivaldo Tro

Chapter 11
Gases

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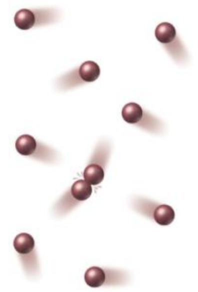
11.2 Kinetic Molecular Theory



1 Collection of particles in constant motion

2 No attractions or repulsions between particles; collisions like billiard ball collisions

3 A lot of space between the particles compared to the size of the particles themselves

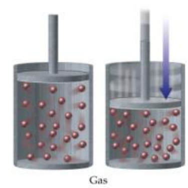
4 The speed that the particles move increases with increasing temperature

Properties

1. **Indefinite Volume**
2. **Indefinite Shape**
   - Rapidly moving particles that do not interact

Gas Properties Explained

3. **Gases are compressible**
4. **Low density**
   - because of the large spaces between the molecules



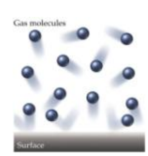
Properties of Gases

5. **Mixtures of gases are always homogeneous**
6. **Gases are fluid**

11.3 Pressure

1. When gas molecules strike a surface, they push on that surface = force
2. The total amount of force is the **pressure** the gas is exerting

pressure = force per unit area

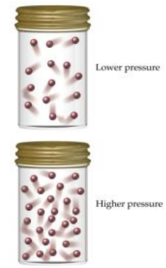


3. Gases will flow from area of high pressure to low pressure
   - ➢ the bigger the difference in pressure, the stronger the flow of the gas
4. If there is something in the gas' path, the gas will try to push it along as the gas flows

The Pressure of a Gas

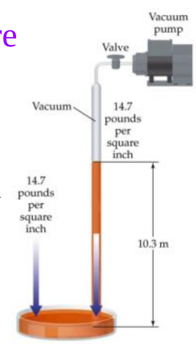
The pressure (P) of a gas depends on several factors:

1. number of gas particles in a given volume = mol
2. volume of the container = V
3. average speed of the gas particles = T (temperature)



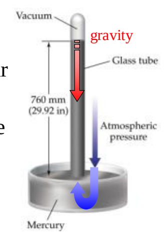
Air Pressure

- The atmosphere exerts a pressure on everything it contacts
  - ✓on average 14.7 psi at sea level
- This pressure holds a column of water about 10.3 m high



Measuring Air Pressure

- use a **barometer**
- column of mercury supported by air pressure
- force of the air on the surface of the mercury is balanced by the pull of gravity on the column of mercury
- Sea level pressure = 760 mm Hg



Common Units of Pressure

Unit	Average Air Pressure at Sea Level
pascal (Pa)	101,325
kilopascal (kPa)	101.325
atmosphere (atm)	1 (exactly)
millimeters of mercury (mm Hg)	760 (exactly)
inches of mercury (in Hg)	29.92
torr (torr)	760 (exactly)
pounds per square inch (psi, lbs./in^2)	14.7

Atmospheric Pressure & Altitude

- The higher up in the atmosphere, the lower the atmospheric pressure.
  - ➢ at sea level atmospheric pressure is 14.7 psi
  - ➢at 10,000 ft it is only 10.0 psi

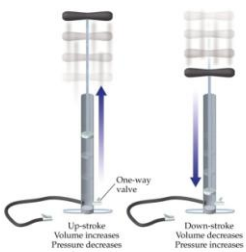
Example 11.1: Converting Between Pressure Units

Example:

- A high-performance road bicycle is inflated to a pressure of 125 psi.
  - ✓ What is the pressure in millimeters of mercury?
  - ✓ What is the pressure in inches of mercury?
  - ✓ What is the pressure in atmospheres?
  - ✓ What is the pressure in pascals?

11.4 Pressure, Volume and Temperature

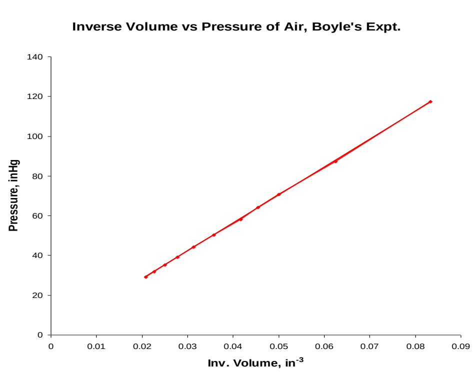
- For a cylinder with a piston (constant T):
  - ✓ Increasing the volume decreases the pressure
  - ✓ Increasing the pressure decreases the volume



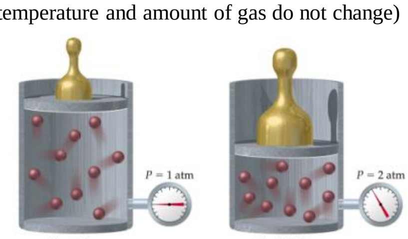
11.4 Boyle's Law

- pressure of a gas is inversely proportional to its volume at constant T and amount of gas.
  - ✓ graph P vs 1/V is straight line
  - ✓ as P increases, V decreases by the same factor
  - ✓ P x V = constant

$$P_1 V_1 = P_2 V_2$$



When you double the pressure on a gas, the volume is cut in half, (as long as the temperature and amount of gas do not change)

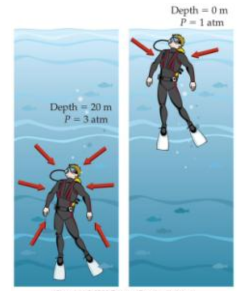


Boyle's Law & Breathing (demo)

- inhale
  - ✓ diaphragm & rib muscles contract
  - ✓ chest cavity expands - volume increase
  - ✓ pressure inside lungs drops below air pressure
  - ✓ air flows into lung to equilibrate pressure
    - ➢ gases move from hi pressure to low
- exhale
  - ✓ diaphragm & rib muscles relax
  - ✓ chest cavity volume decreases
  - ✓ pressure inside lungs rises above air pressure
  - ✓ air flows out of lung to equilibrate pressure
- normal healthy person can generate a lung pressure of 1.06 atm

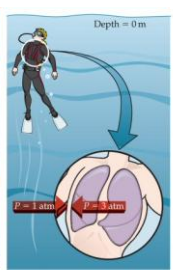
Boyle's Law and Diving

- In water, for each 10 m you dive below the surface the pressure on your lungs increases 1 atm
  - ✓ at 20 m the total pressure is 3 atm
  - ✓ if your tank contained air at 1 atm pressure you would not be able to inhale it into your lungs

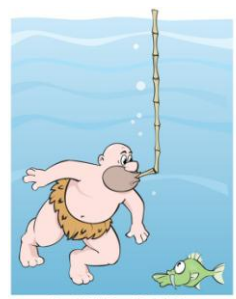


Boyle's Law and Diving

- scuba tanks have a regulator so that the air in the tank is delivered at the same pressure as the water surrounding you
- if a diver holds her breath and rises quickly, so that the outside pressure drops to 1 atm; according to Boyle's Law, what should happen to the volume of air in the lungs?



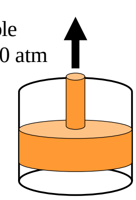
Is this possible at a depth of 20 m?



Example 11.2: Boyle's Law

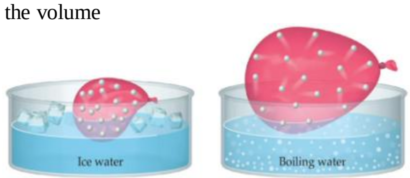
Example:

- A cylinder equipped with a moveable piston has an applied pressure of 4.0 atm and a volume of 6.0 L. What is the volume if the applied pressure is decreased to 1.0 atm?



11.5 Pressure, Volume and Temperature

- For a balloon the pressure outside and inside is the same (<u>constant pressure</u>):
  - ✓Decreasing the temperature causes the balloon to decrease its volume.
  - ✓Raising the temperature causes the balloon to increase its the volume

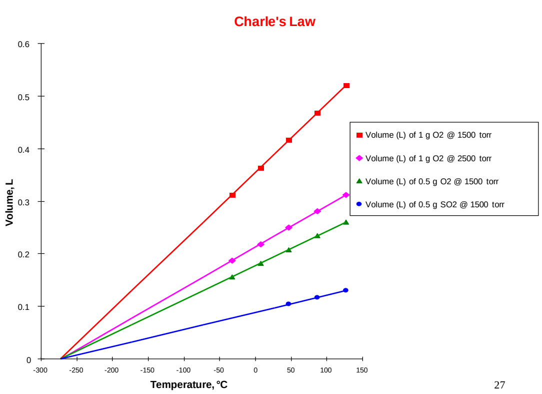


<u>Charles' Law</u>

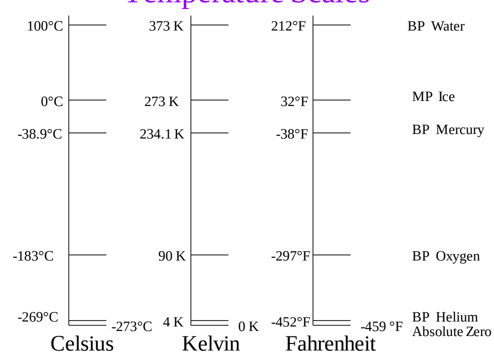
volume is directly proportional to temperature at constant P and amount of gas

- graph of V vs T is straight line
- as T increases, V also increases
- V = constant x T
  (if T measured in Kelvin, K = °C + 273)

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



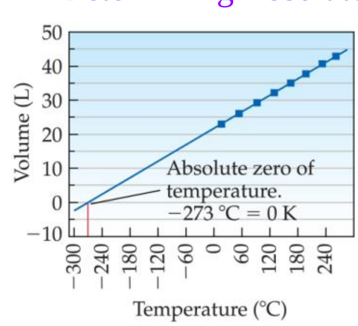
Temperature Scales



Absolute Zero

- theoretical ***temperature*** at which a gas would have zero volume and no pressure
  - ✓Kelvin calculated by extrapolation
- 0 K = -273.15 °C = -459 °F = 0 R
- never attainable
  - ✓though we've gotten real close!
- all gas law problems use the Kelvin temperature scale!

Determining Absolute Zero

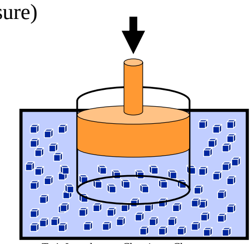


William Thomson, the Lord of Kelvin, extrapolated the line graphs of Volume vs. Temp. to determine the theoretical temp. a gas would have a volume of 0.

Example 11.3: Charles' Law

Example:

- A sample of gas has a volume of 2.80 L at an unknown temperature. When the sample is submerged in ice water at 0°C, its volume decreases to 2.57 L. What was the initial temperature in kelvin and in celsius? (assume constant pressure)



Standard Conditions

- Common reference points for comparing
- standard pressure = 1.00 atm
- standard temperature = 0°C = 273 K
- STP

11.6 The Combined Gas Law

- Boyle's Law shows the relationship between pressure and volume
  - ✓ at constant temperature
- Charles' Law shows the relationship between volume and absolute temperature
  - ✓ at constant pressure
- the two laws can be combined together to give a law that predicts what happens to the volume of a sample of gas when both the pressure and temperature change

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

Example 11.4: The Combined Gas Law

Example 1:

- A sample of gas has an initial volume of 158 mL at a pressure of 735 mmHg and a temperature of 34°C. If the gas is compressed to a volume of 108 mL and heated to 85°C, what is the final pressure in mmHg?

Example 2:

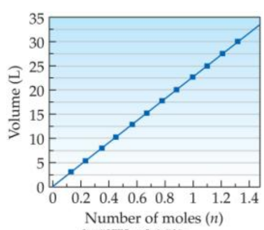
- A sample of gas has an initial volume of 158 mL at a pressure of 735 mmHg and a temperature of 34°C. What will be its volume at STP?

11.7 Avogadro's Law

- volume directly proportional to the number of gas molecules
  - ✓**V** = constant x **n**
  - ✓constant P and T
  - ✓more gas molecules = larger volume
- count number of gas molecules by **moles**
- equal volumes of gases contain equal numbers of molecules
  - ✓***the gas doesn't matter***

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

Avogadro's Law



Example 11.5: Avogadro's Law

Example:

- A 4.8 L sample of helium gas contains 0.22 mol helium. How many additional moles of helium must be added to obtain a volume of 6.4 L? (assume constant pressure and temperature)

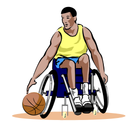
11.8 Ideal Gas Law

- By combing 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
  - ✓ we will use $0.0821 \frac{atm \bullet L}{mol \bullet K}$ and convert P to atm and V to L
- use the Ideal Gas law when have a gas at one condition, use the Combined Gas Law when you have gas whose condition is changing

$$\frac{(P)(V)}{(n)(T)} = R \quad \text{or} \quad PV = nRT$$

Example 11.7: The Ideal Gas Law Requiring Unit Conversion

Example:
- Calculate the number of moles of gas in a basketball inflated to a total pressure of 24.2 psi with a volume of 3.2 L at 25°C



Molar Mass of a Gas

- one of the methods chemists use to determine the molar mass of an unknown substance is to heat a weighed sample until it becomes a gas, measure the temperature, pressure and volume, and use the Ideal Gas Law

$$\text{Molar Mass} = \frac{\text{mass in grams}}{\text{moles}}$$

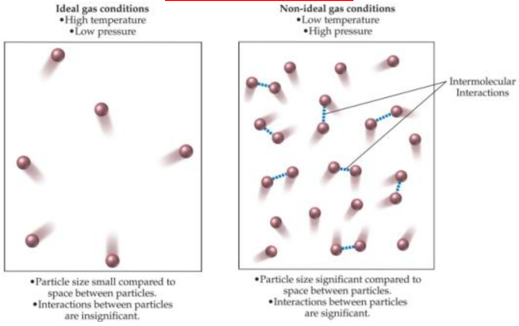
Example 11.8: Molar Mass Using The Ideal Gas Law and a Mass Measurement

Example:
- A sample of a gas has a mass of 0.311 g. Its volume is 0.225 L at a temperature of 55°C and a pressure of 886 mmHg. Find its molar mass.

Ideal vs. Real Gases

- Real gases often do not behave like ideal gases at high pressure or low temperature
- Ideal gas laws assume
  1) no attractions between gas molecules
  2) gas molecules do not take up space
  3) based on the Kinetic-Molecular Theory
- at low temperatures and high pressures these assumptions are not valid

Ideal vs. Real



Tro's Introductory Chemistry, Chapter 49

11.9 Mixtures of Gases

- According to Kinetic Molecular Theory, the particles in a gas behave independently
- Air is a mixture, yet we can treat it as a single gas
- Also, we can think of each gas in the mixture independent of the other gases
  - ✓ though all gases in the mixture have the same volume and temperature
    - ➢ all gases completely occupy the container, so all gases in the mixture have the volume of the container

Gas	% in Air, by volume	Gas	% in Air, by volume
nitrogen, N_2	78	argon, Ar	78
oxygen, O_2	21	carbon dioxide, CO_2	21

50

Partial Pressure

- each gas in the mixture exerts a pressure independent of the other gases in the mixture
- the pressure of an component gas in a mixture is called a **partial pressure**
- the sum of the partial pressures of all the gases in a mixture equals the total pressure
  - ✓ Dalton's Law of Partial Pressures
  - ✓ $P_{total} = P_{gas\ A} + P_{gas\ B} + P_{gas\ C} + \ldots$

$P_{air} = P_{N_2} + P_{O_2} + P_{Ar} = 0.78\ atm + 0.21\ atm + 0.01\ atm = 1.00\ atm$

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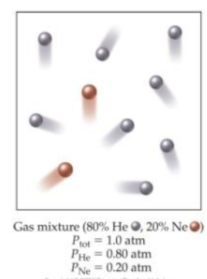
Finding Partial Pressure

- to find the partial pressure of a gas, multiply the total pressure of the mixture by the fractional composition of the gas
- for example, in a gas mixture that is 80.0% He and 20.0% Ne that has a total pressure of 1.0 atm, the partial pressure of He would be:

$P_{He} = (0.800)(1.0\ atm) = 0.80\ atm$

  - ✓ fractional composition = percentage divided by 100



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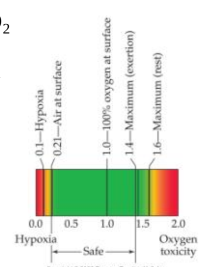
Mountain Climbing & Partial Pressure

- our bodies are adapted to breathe O_2 at a partial pressure of 0.21 atm
  - ✓ Sherpa, people native to the Himalaya mountains, are adapted to the much lower partial pressure of oxygen in their air
- partial pressures of O_2 lower than 0.1 atm will lead to **hypoxia**
  - ✓ unconsciousness or death
- climbers of Mt Everest must carry O_2 in cylinders to prevent hypoxia
  - ✓ on top of Mt Everest, P_{air} = 0.311 atm, so P_{O2} = 0.065 atm



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Deep Sea Divers & Partial Pressure

- its also possible to have too much O_2, a condition called **oxygen toxicity**
  - ✓ P_{O2} > 1.4 atm
  - ✓ oxygen toxicity can lead to muscle spasms, tunnel vision and convulsions
- its also possible to have too much N_2, a condition called **nitrogen narcosis**
  - ✓ also known as Rapture of the Deep
- when diving deep, the pressure of the air divers breathe increases – so the partial pressure of the oxygen increases
  - ✓ at a depth of 55 m the partial pressure of O_2 is 1.4 atm
  - ✓ divers that go below 50 m use a mixture of He and O_2 called heliox that contains a lower percentage of O_2 than air

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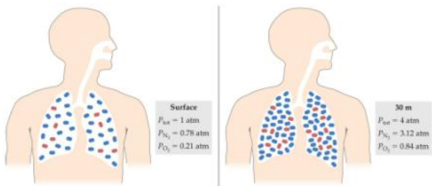
Partial Pressure vs. Total Pressure

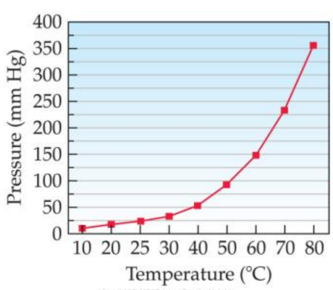


At a depth of 30 m, the total pressure of air in the divers lungs, and the partial pressure of all the gases in the air, are quadrupled!

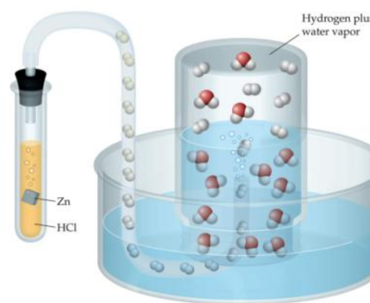
Collecting Gases

- gases are often collected by having them displace water from a container
- the problem is that since water evaporates, there is also water vapor in the collected gas
- the partial pressure of the water vapor, called the **vapor pressure**, depends only on the temperature
  - ✓ so you can use a table to find out the partial pressure of the water vapor in the gas you collect
- if you collect a gas sample with a total pressure of 758 mmHg at 25°C, the partial pressure of the water vapor will be 23.8 mmHg – so the partial pressure of the dry gas will be 734 mmHg

Vapor Pressure of Water



Temp., °C	**Pressure, mmHg**
10	9.2
20	17.5
25	23.8
30	31.8
40	55.3
50	92.5
60	149.4
70	233.7
80	355.1



Zn metal reacts with HCl(aq) to produce H_2(g).

The gas flows through the tube and bubbles into the jar, where it displaces the water in the jar.

Because water evaporates, some water vapor gets mixed in with the H_2.

11.10 Reactions Involving Gases

- the principles of reaction stoichiometry from Chapter 8 can be combined with the Gas Laws for reactions involving gases
- in reactions of gases, the amount of a gas is often given as a Volume
  - ✓ instead of moles
  - ✓ as we've seen, must state pressure and temperature
- the Ideal Gas Law allows us to convert from the volume of the gas to moles; then we can use the coefficients in the equation as a mole ratio

Example 11.11: Gases in Chemical Reactions

Example:

- How many liters of oxygen gas form when 294 g of $KClO_3$ completely reacts in the following reaction? Assume the oxygen gas is collected at P = 755 mmHg and T = 308 K

$$2\,KClO_3(s) \xrightarrow{\Delta} 2\,KCl(s) + 3\,O_2(g)$$

Calculate the volume occupied by 1.00 moles of an ideal gas at STP.

P x V = n x R x T

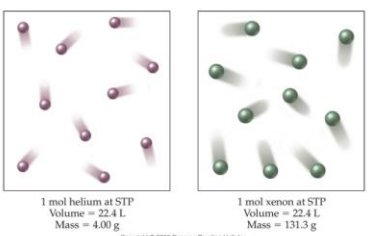
$$(1.00\text{ atm}) \times V = (1.00\text{ moles})(0.0821\,\tfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(273\text{ K})$$

V = 22.4 L

- 1 mole of any gas at STP will occupy 22.4 L
- this volume is called the **molar volume** and can be used as a conversion factor
  - ✓ as long as you work at STP

1 mol ≡ 22.4 L

Molar Volume



There is so much empty space between molecules in the gas state, the volume of the gas is not effected by the size of the molecules, (under ideal conditions).

Example 11.12: Using Molar Volume in Calculations

Example:

- How many grams of water will form when 1.24 L of H_2 at STP completely reacts with O_2?

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$