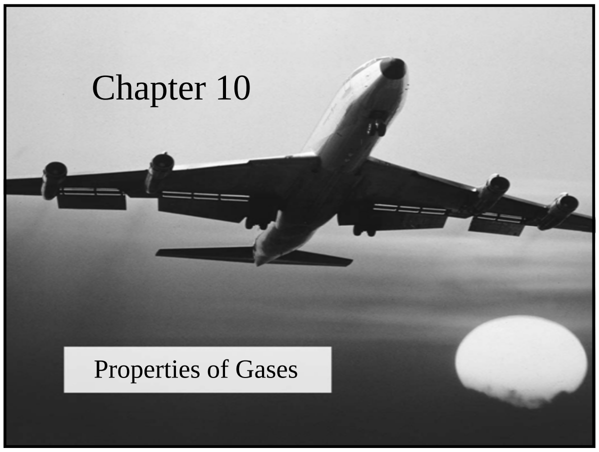


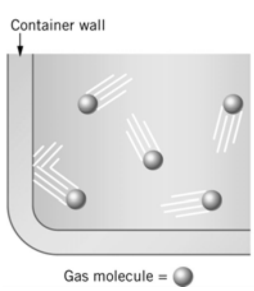
10.1 Familiar properties of gases

1. Gases are compressible.
2. Gases exert a pressure.
3. Pressure of the gases depends on how much gas is confined.
4. Gases fill completely any container into which they are placed.
5. Gases mix freely and quickly with each other.
6. The pressure of a gas rises when its temperature is increased.

2

Molecular model

- These properties suggest a molecular model
  - Lots of space between the molecules
  - Moving rapidly



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10.2 Pressure

Pressure (P) – the force which acts on a given area

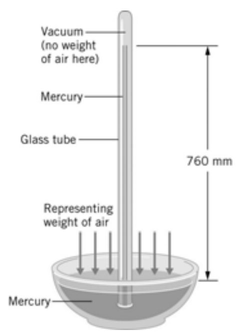
$$P = \frac{F}{A}$$

http://wine1.sb.fsu.edu/chm1045/chm1045.htm

4

Atmospheric Pressure

- Pressure (P) – in a gas it is caused by molecular collisions with a container.
- The air around us is a huge reservoir of gas that exerts pressure on the Earth's Surface
- The atmospheric pressure can be measured with a mercury barometer.
- Why does pressure have units of mmHg?



Pressure Units

- There are several units that are used for pressure in addition to mmHg.

  1 mmHg = 1 torr
- Because the pressure at sea level is 760 mm Hg we define another unit called the atmosphere

  1 atm = 760 mmHg (exactly)
- The official SI unit is the pascal

  1 atm = 101325 pascal = 101325 Pa

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Open-end Manometer

- How do we measure the pressure of a gas in a closed container?
- We must use an <u>open-end manometer</u>.
- In the U-shaped manometer, the difference between the two mercury levels gives the pressure in mmHg of the gas.

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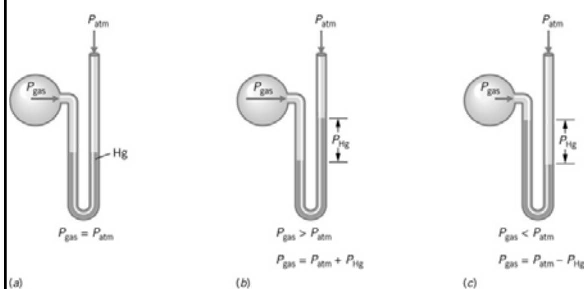
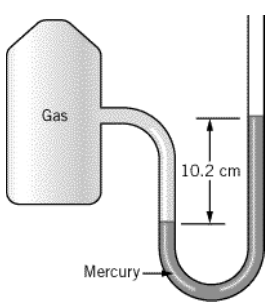
Open-end Manometer



8

A student collected a gas in an apparatus connected to an open-end manometer, as illustrated in the figure. The difference in heights of mercury in the two columns was 102 mm and the atmospheric pressure was measured to be 756 mmHg. What was the pressure of the gas in the apparatus in atm?



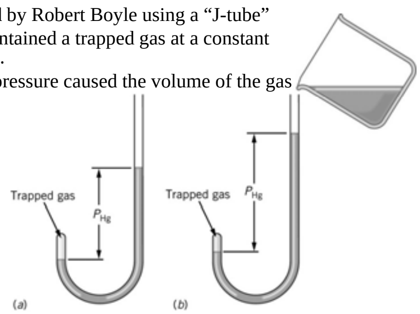
10.3 The Gas Laws

- There are four variables that affect the properties of a gas:
  - Pressure
  - Volume
  - Kelvin Temperature
  - Amount of the gas
- Simple experiments can be conducted that relate how these variables change
- The gas laws summarize these experiments.

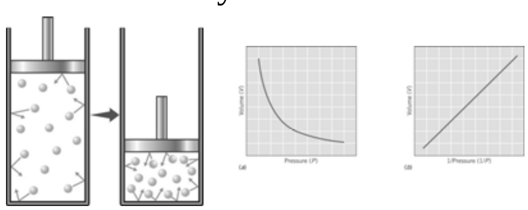
10

Volume and Pressure

- Investigated by Robert Boyle using a "J-tube"
- The tube contained a trapped gas at a constant temperature.
- Increasing pressure caused the volume of the gas to decrease.



Boyle's Law



- Volume is inversely proportional to pressure.

$$V_{gas} \propto \frac{1}{P_{gas}} \quad \text{at fixed temperature and fixed amount}$$

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Ideal Gas

- If you take very precise measurements, Boyle's law doesn't quite work.
- No gas obeys the law over a wide range of temperatures and pressures.
- To make things easier on us, we will study hypothetical gases.
- **Ideal gas** – a gas that obeys the gas laws over all temperatures and pressures.

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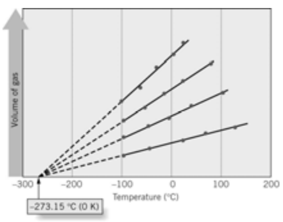
Charles' Law

- Jacques Charles reported studies of gas volume as a function of temperature.
- He determined that the volume of gas is directly proportional to its **Kelvin** temperature.



$V_{gas} \propto T_{gas}$ at fixed pressure and fixed amount

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Gay-Lussac's Law

- The pressure of a fixed amount of gas held at a constant volume is directly proportional to the Kelvin temperature.
- This would happen when a gas is in a closed container.

$P_{gas} \propto T_{gas}$ at fixed volume and fixed amount

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Combined Gas Law

- A combination of Boyle’s, Charles’s and Gay-Lussac’s work
- It only applies to situations where the *quantity (how many moles) of the gas does not change.*

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

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A sample of nitrogen has a volume of 883 mL and a pressure of 741 torr. What pressure will change the volume to 655 mL at the same temperature?

17

What will be the final pressure of a sample of nitrogen with a volume of 955 mL at 745 torr and 25˚C if it is heated to 62 ˚C and given a final volume of 1155 mL?

18

10.4 (Avogadro's Principle)

- When measured at the same temperature and pressure, equal volume of gases contain equal numbers of moles.
- Gas volume is directly proportional to the amount of gas.

$V_{gas} \propto n_{gas}$ at fixed pressure and fixed temperature

- What does this mean for us? (think stoichiometry)

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Some Definitions

- STP = standard temperature and pressure
  - T = 273 K and P = 1 atm
- Standard Molar Volume = the volume that one mole of a gas occupies at STP
  - V = 22.414 L

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Stoichiometry at STP

- Gas volume can be used in solving stoichiometry problems, provided they are measured at the same temperature and pressure.
- Problem: How many liters of hydrogen, measured at STP, are needed to combine exactly with 1.50 L of nitrogen, also at STP, to form ammonia?

$$N_2 + 3\ H_2 \rightarrow 2\ NH_3$$

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10.5 The Ideal Gas Law

- All four variables (P, T, V, and n) can be related using a single constant.
- The universal gas constant, R.

$$PV = nRT$$

- The gas constant: R = 0.0821 L atm/mol K
- Units: n (moles), P (atm), V (L) and T (K)

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Steps for doing gas problems

1. Write down what you know (convert temperature to K).
2. Write down what you are looking for.
3. Determine what equation you will be using.
4. Convert pressure to atm and volume to L if using ideal gas law.
5. Complete calculation.
6. Check and make sure answer is in the correct units, if not → convert!

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How many grams of Ar are in a 12.0 L cylinder of argon at a pressure of 57.8 atm and a temperature of 25˚C?

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Molar Mass and the Ideal Gas Law

- You can use the ideal gas law to determine molar mass.
- You only need two pieces of information.
  - The mass (g) of the gas
  - The amount (mol) of the gas.
- Remember the units on molar mass (g/mol)

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The label on a cylinder of a noble gas became illegible, so a student allowed some of the gas to flow into a gas bulb with a volume of 0.3000L until the pressure was 0.901 atm. The sample now weighed 1.45 g and its temperature was 27.0˚C. What is the molar mass of the gas? Which noble gas is it?

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Density and the Ideal Gas Law

- Remember, density?
- What are the units of density?
- For gases, it is usually g/L.
- You can also use the ideal gas law equation to determine density.
- You will need
  - Mass (g)
  - Volume (L)

27

One procedure used to separate the isotopes of uranium to obtain material to construct a nuclear weapon employs uranium compound with the formula UF_6. The compound boils at about 56˚C, so at 100.˚C it is a gas. What is the density of 1.00 mol of UF_6 at 100.˚C if the pressure of the gas is 0.974 atm. (MM of UF_6 = 352.0 g/mol)

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Gas Stoichiometry

- Stoichiometry applies to solids, liquids and gases
- Notice that the ideal gas law contains moles.
- And remember that in order to use stoichiometry, you must be in units of moles.

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Suppose 1.25 g sample of $CaCO_3$ is decomposed by heating. How many milliliters of CO_2 gas will be evolved if it will be measured at 0.974 atm and 25˚C? (MM of $CaCO_3$ = 100.05 g/mol)

$$CaCO_3\ (s) \rightarrow CO_2\ (g) + CaO\ (s)$$

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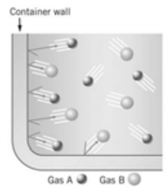
10.5 Partial Pressures



- In a mixture, each gas contributes to the total pressure.
- <u>Partial pressure</u> – the amount of pressure due to a specific gas in a mixture.
- If a mixture contains gases A, B and C, then the partial pressure A is P_A. What is the partial pressure of B and C?

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Dalton's Law of Partial Pressures

- The total pressure of a mixture of gases is the sum of their individual partial pressures.

$$P_{total} = P_A + P_B + P_C + \cdots$$

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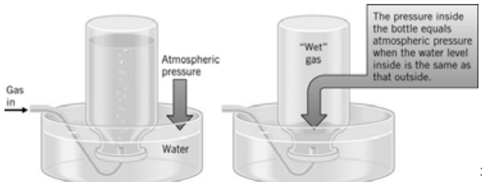
Gas Collection Over Water

- Many chemistry experiments require the collection of gas over water.
- What is this?



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Gas Collection Over Water

- The space above any liquid contains some of the liquid's vapor
- The pressure this vapor exerts is called the vapor pressure.
- As the gas bubbles through the water, water vapor gets into the gas so the total pressure inside the bottle includes the partial pressure of the water vapor and the partial pressure of the gas.
- You must subtract off the P_{H_2O} to obtain the pressure of the gas.
- You can look up the P_{H_2O} for specific temperatures.

$$P_{gas} = P_{total} - P_{H_2O}$$

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TABLE 10.2 Vapor Pressure of Water at Various Temperatures

Temperature (°C)	Vapor Pressure (torr)	Temperature (°C)	Vapor Pressure (torr)
0	4.579	50	92.51
5	6.543	55	118.0
10	9.209	60	149.4
15	12.79	65	187.5
20	17.54	70	233.7
25	23.76	75	289.1
30	31.82	80	355.1
35	41.18	85	433.6
37[a]	47.07	90	525.8
40	55.32	95	633.9
45	71.88	100	760.0

[a] Human body temperature.

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A sample of oxygen is collected over water at 20˚C and a pressure of 738 torr. What is the partial pressure of oxygen in mmHg?

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Gas mixtures continued…

- The composition of a gas mixture can be expressed in mole fractions and mole percents.
- Mole fraction – the ratio of the number of moles of a given component to the total number of moles of all components.

$$X_A = \frac{n_A}{n_A + n_B + n_c + \cdots} = \frac{n_A}{n_{total}}$$

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Mole Percent

Mole percent (mol%) = mole fraction * 100%

Practice Problem: What is the mole fraction (and mole percent) of each component of a mixture composed of 0.200 mol O_2 and 0.500 mol N_2?

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Mole Fraction and Partial Pressures

- Mole fractions and partial pressures are related.

$$X_A = \frac{P_A}{P_{total}}$$

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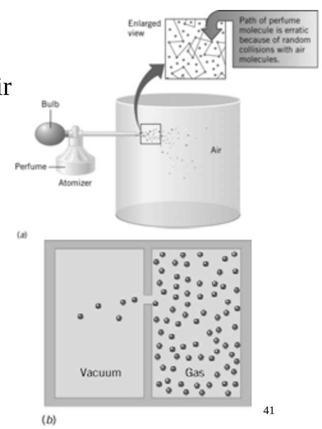
Problem: What is the mole fraction and the mole percent of oxygen in exhaled air if the P_{O_2} = 760. torr and P_{total} is 1116 torr?

40

Diffusion – mixing of two gases due to their molecular motion

Effusion – movement of gas through a tiny opening into a vacuum



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10.7 Graham's Law

- Obviously, the heavier the gas molecule, the slower it will diffuse or effuse.
- We can relate the molar mass of a molecule to effusion (and diffusion).

$$\frac{\text{effusion rate of A}}{\text{effusion rate of B}} = \sqrt{\frac{MM_B}{MM_A}}$$

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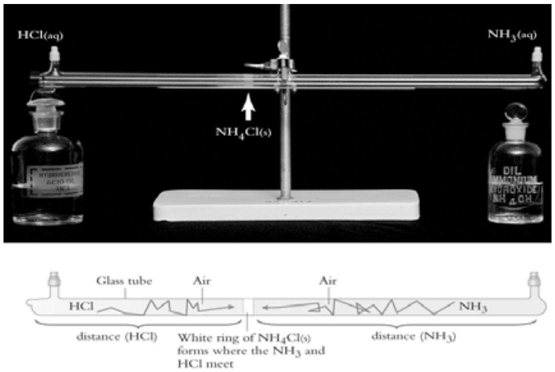

WHY?

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Rank the follow gases in order of increasing rates of effusion: O_2, Ar, CO_2, N_2, He, and SO_2

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10.8 Kinetic Molecular Theory Explains Gas Behaviors

- Gases consists of an extremely large number of very tiny particles that:
  - are in constant, random motion
  - occupy a negligible portion of the total volume of the sample-their individual contribution may be ignored
  - collide elastically with themselves and the walls of the container
  - move in straight lines between collisions, neither attracting nor repelling each other

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Kinetic Molecular Theory- Irregularities

- The volume of a gas molecule is negligible
  - NO! Under conditions of extremely high pressure, gases are closer, their relative size is a factor
- Gas molecules collide elastically
  - NO! Under conditions of extremely low temperatures, gases move more slowly and intermolecular attractions are more significant

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10.9 Real Gases (nonideal)

- During conditions of high pressure and low temperature and when forces exist between molecules
  - At high pressures, molecular volume becomes important
  - When collisions are inelastic, the forces between the molecules become important

$$\left[P + a\left(\frac{n}{V}\right)^2\right]\left[V - bn\right] = nRT$$

Van der Waals Equation, corrects for these problems

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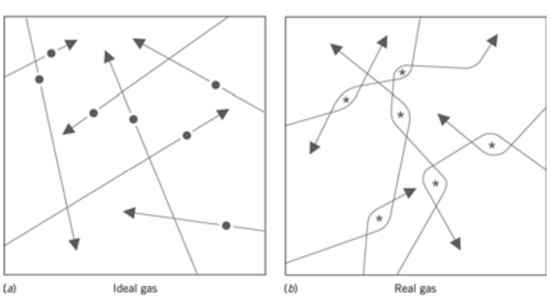

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