States of Matter

CHAPTER 5 GASES
CHAPTER 10 LIQUIDS AND SOLIDS
CHAPTER 11 SOLUTIONS
CHAPTER 6 ORGANIC CHEMISTRY

Homework

- AP Test Chapter 5
- AP Test Chapter 10
- AP Test Chapter 11

Equation sheet

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{total} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

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

$$n = \frac{m}{\boldsymbol{M}}$$

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

$$D = \frac{m}{V}$$

$$KE_{molecule} = \frac{1}{2}mv^2$$

Molarity, M = moles of solute per liter of solution

$$A = \varepsilon bc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
$\boldsymbol{M}$ = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
ε = molar absorptivity
b = path length
c = concentration

Gas constant, R = 8.314 J mol^{-1} K^{-1}
= 0.08206 L atm mol^{-1} K^{-1}
= 62.36 L torr mol^{-1} K^{-1}

1 atm = 760 mm Hg = 760 torr

STP = 273.15 K and 1.0 atm

Ideal gas at STP = 22.4 L mol^{-1}

Pressure

- A **barometer** is a device to measure atmospheric pressure.
- The barometer was invented in 1643 by a student of Galileo, Italian scientist Evangelista Torricelli.
- His barometer is constructed by filling a glass tube with liquid mercury and inverting it in a dish of mercury.
- Mercury is used to measure pressure because of its high density. By way of comparison, the column of water required to measure a given pressure would be 13.5 times as high as a mercury column used for the same purpose.

Barometer

- Device used to measure atmospheric pressure.
- Mercury flows out of the tube until the pressure of the column of mercury standing on the surface of the mercury in the dish is *equal* to the pressure of the air on the rest of the surface of the mercury in the dish.

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- Atmospheric pressure results from the mass of the air being pulled toward the center of the Earth by gravity.
- Altitude also affects the pressure.
- * At sea level, standard pressure is 760 mm of mercury, 101 kPa or 1.00 atm.
- * At the peak of Mt. Everest, average pressure 270 mm Hg, 36 kPa, or .356 atm

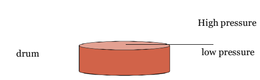
Weather

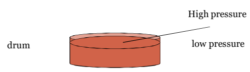
- Different pressures correspond to different weather.
- Low pressure is stormy
- High pressure is clear
- Moist air rises, condenses, as it cools, forms clouds… Air sinks, it warms as it falls

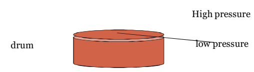
Low High

Modern Barometers

- Digital Barometers and barometers with a dial use a sensor on a sealed drum.
- The top of the drum is flexible.
- Sealed inside the drum is air at a known (calibrated) pressure.
- Higher outside pressure caves the drum in.
- Lower outside pressure bows the drum out.

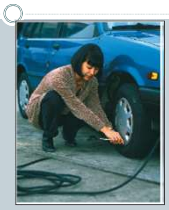






Manometer

- Device used for measuring the pressure of a gas in a container.



Common Units of Pressure.

- 1 atmosphere (atm) = 760 mm Hg/torr
- 101.325 kPa

- Pressure Conversions.
- The pressure of a gas is measured as 49 torr. Represent this pressure in both atmospheres and kilopascals.

Boyle's Law

- Named after Irish chemist Robert Boyle (1627-1691).
- Using a closed-ended J-tube, Boyle found that gas volume is inversely proportional to its pressure.
- With technology we have found that Boyle's law holds precisely only at very low pressures.

- If a gas obey Boyle's law exactly, it is said to be "ideal."
- Boyle's Law: $PV = k$ or $P_1V_1 = P_2V_2$.

- Sulfur dioxide (SO_2), a gas that plays a central role in the formation of acid rain, is found in the exhaust of automobiles and power plants. Consider a 1.53-L sample of gaseous SO_2 at a pressure of 5.6 kPa. If the pressure is changed to 15 kPa at a constant temperature, what will be the new volume of the gas?

Charles' Law

- Named in honor of Jacques Charles (1746-1823).
- He was on the first manned hydrogen balloon flight.
- He found that the volume of a gas at constant pressure increases linearly with the temperature of the gas.
- The temperature was originally plotted in degrees *C*.
- A plot of the gases showed that all volumes extrapolated to the same temperature, -273.2 *C*.
- This temperature is also known as 0 Kelvin or *absolute zero*.
- The Kelvin scale was devised by English physicist William Thomson, also known as Lord Kelvin, 50 years after Charles's conclusions.

Charles' Law

- Charles's Law: $V/T = k$ or $V_1/T_1 = V_2/T_2$
- * Temperature is in *Kelvins*.

- A sample of gas at 15° *C* and 1 atm has a volume of 2.58 L. What volume will the gas occupy at 38° *C* and 1 atm?

Avogadro's Law

- Named in honor of Italian chemist Amadeo Avogadro (1776-1856).
- He stated that at the same temperature and pressure, equal volumes of different gases contain the same number of particles.

- Avogadro's Law: $V/n = k$ or $V_1/n_1 = V_2/n_2$

- At 1 atm of pressure and 273 K, 1 mol of any gas occupies 22.4 L.

Avogadro's Law

- Suppose we have a 12.2 L sample containing 0.50 mol oxygen gas (O_2) at a pressure of 1.0 atm and a temperature of 25° C. If all this O_2 is converted to ozone (O_3) at the same temperature and pressure, what would be the volume of the ozone?

Combined/Ideal Gas Law

- Combining the previous laws gives us the combined gas law
- VP/T = VP/T
- A common derivation of the combined gas law is the ideal gas law
- PV = nRT
- *R* is the **universal gas constant**.
- These are given on equation sheet
- R = **8.314** J/mol K or L kPa/mol K
- **R** = **0.08206** L atm/mol K
- R = **62.36** L mm Hg/mol K

Molecular View of The Ideal Gas Law

Gas Laws I

- A sample of hydrogen gas (H_2) has a volume of 8.56 L at a temperature of 0° C and a pressure of 1.5 atm. Calculate the moles of H_2 molecules present in this sample of gas.

- Suppose we have a sample of ammonia gas with a volume of 3.5 L at a pressure of 1.68 atm. The gas is compressed to a volume of 1.35 L at a constant temperature. Calculate the final pressure.

Gas Laws II

- A sample of methane gas that has a volume of 3.8 L at 5° C is heated to 86° C at constant pressure. Calculate its new volume.
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- A sample of diborane gas (B_2H_6), a substance that bursts into flames when exposed to air, has a pressure of 345 torr at a temperature of -15° C and a volume of 3.48 L. If conditions are changed so that the temperature is 36° C and the pressure is 468 torr, what will be the volume of the sample?

Gas Laws III

- A sample containing 0.35 mol argon gas at a temperature of 13° C and a pressure of 568 torr is heated to 56° C and a pressure of 897 torr. Calculate the change in volume that occurs.

Gas Stoichiometry

- **STP** stands for **standard temperature and pressure**.
- The temperature = 0° C = 273 K.
- NOT TO BE CONFUSED WITH STANDARD STATE 25° C or 298K!!!

- The pressure is 1.00 atm, 101 kPa or an equivalent.
- **Molar volume** under these condition for an ideal gas is 22.4 L.
- All of this is on the equation sheet

Gas Stoichiometry

- A sample of nitrogen gas has a volume of 1.75 L at STP. How many moles of N_2 are present?

- Quicklime (CaO) is produced by the thermal decomposition of calcium carbonate ($CaCO_3$). Calculate the volume of CO_2 at STP produced from the decomposition of 152 g $CaCO_3$ by the reaction
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- $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

Gas Stoichiometry II

- A sample of methane gas having a volume of 2.80 L at 25° C and 1.65 atm was mixed with a sample of oxygen gas having a volume of 35.0 L at 31° C and 1.25 atm. The mixture was then ignited to form carbon dioxide and water. Calculate the volume of CO_2 formed at a pressure of 2.50 atm and a temperature of 125° C.

Molar Mass of a Gas and Gas Density.

- Deriving the equations for molar mass of a gas and gas density:
- n = m/M ---
- m = mass (g) M = molar mass (g/mol) n number of particles (mol)
- PV = nRT   PV = m/MRT
- M = mRT/PV
- d = m/v (D= density)
- M = dRT/P

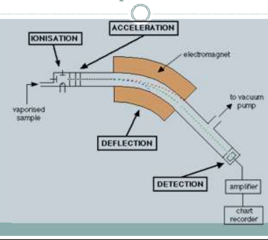
Molar mass Determination

- The determination of molar masses of elements are relative to one another.
- That is to say Avoagadro took a certain amount of a gas and said, "this is a mole". He took the mass of that amount, and that became the molar mass.
- HE THEN STARTED REACTING THE GAS WITH OTHER THINGS TO DETERMINE WHAT A MOLE OF THAT WAS.
- He took the mass of the other elements, and that became the molar mass of that element relative to the first.

Molar mass

- The current periodic table is based off of the most common isotope of carbon, C-12, having a molar mass 12.00 g/mol.
- The best method for determining molar mass today uses a **mass spectrometer**.
- To use this, elements are turned into ions and passed through a magnetic field. This deflects the ions.
- Mass is inertia, resistance to change in motion. Heavier atoms will deflect less. The amount of deflection is measured and used to calculate the atomic mass.

Mass spectrometer



Gas Density/Molar Mass.

- The density of a gas was measured at 1.50 atm and 27° C and found to be 1.95 g/L. Calculate the molar mass of the gas.

Dalton's Law of Partial Pressure

- Named in honor of English scientist John Dalton.
- He stated that gases mix homogeneously.
- He also said that each gas in a mixture behaves as if it were the only gas present.
- Dalton' Law: For a mixture of gases in a container, the total pressure exerted is the sum of the pressures that each gas would exert if it were alone.
- $P_{total} = P_1 + P_2 + P_3 + \ldots$

- The fact that the pressure exerted by an ideal gas is not affected by the identity (composition) of the gas particle reveal two things.
- The volume of the individual gas particles must not be important.
- The forces among the particles must not be important.