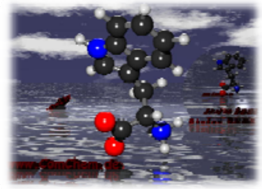
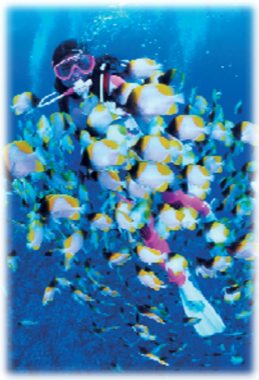
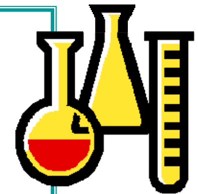




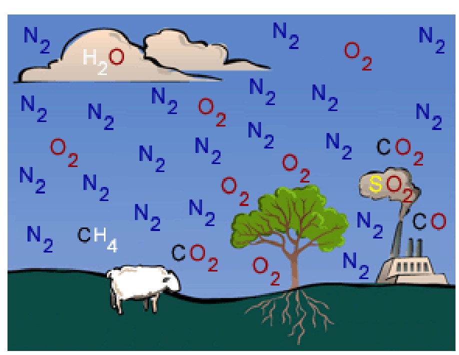
Chapter 11

Gases



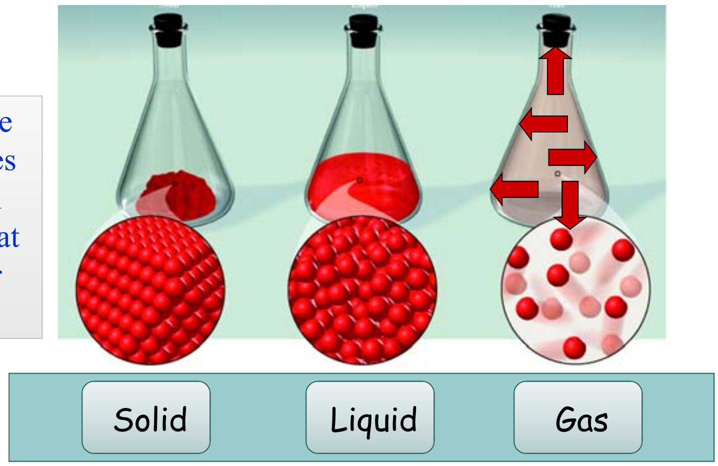
Preview

- Properties and measurements of gases.
- Effects of temperature, pressure and volume.
  - Boyle's law.
  - Charles's law, and
  - Avogadro's law.
- The ideal gas equation.
- Gas mixtures and partial pressure.
- The Kinetic Molecular Theory.
- Real gases.



States of Matter

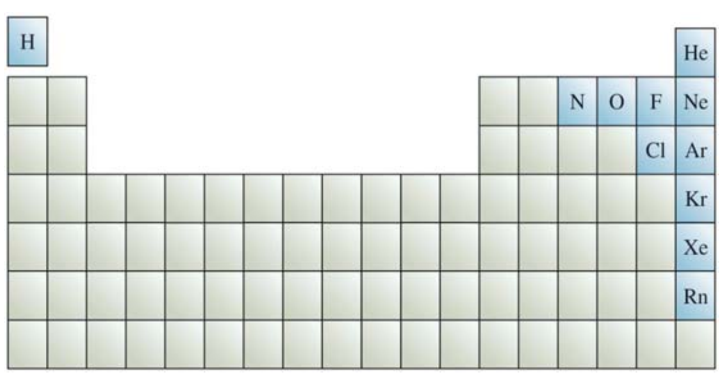
How do the three phases vary from each other at molecular levels?



Unlike liquid and solid, gas fills up the container and exerts force towards the inner walls of the container.

Gaseous Elements

- Few elements and many low molar masses molecular compounds do exist as gases at room temperature.



Chapter 11 Section 1

Gaseous Compounds

- Few elements and many low molar masses molecular compounds do exist as gases at room temperature.

TABLE 11.1 Molecular Compounds That Are Gases at Room Temperature

Molecular Formula	Compound Name
HCl	Hydrogen chloride
NH_3	Ammonia
CO_2	Carbon dioxide
N_2O	Dinitrogen monoxide or nitrous oxide
CH_4	Methane
HCN	Hydrogen cyanide

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Chapter 11 Section 1

Characteristics of Gases

- The gas is different from condensed phases (liquid and solid) in many ways:
  1. It assumes both the shape and volume of its container.
  2. It is compressible.
  3. It has a lower density compared to liquid and solid.
  4. It forms homogeneous mixtures with other gases (that it doesn't react with) in any proportions "miscibility".

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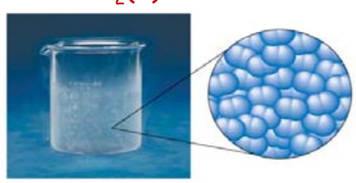
States of Gas and Liquid for Nitrogen

1 mol $N_2(l)$



1 mol $N_2(g)$

	1 mol $N_2(l)$	1 mol $N_2(g)$
Volume	35 mL	22.4 L
Density	0.81 g/mL	0.0012 g/mL
Temperature	~ -190°C	Room temperature

Liquid nitrogen freezes at -210°C and boils at -195°C !!

← Liquid nitrogen boiling in a cup

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Chapter 11 Section 1

How Powerful the Gas is?

- A sample of a gas confined to a container exerts a ***pressure*** on its walls. An example is the tire of a car.

An empty metal can

Air inside the can is removed

Atmospheric pressure crashes the can

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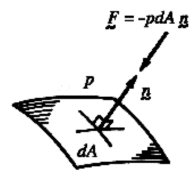
Chapter 11 Section 1

Definition of Pressure

- ***Pressure*** is defined as the force applied per unit area
- $P = \dfrac{\text{Force}}{\text{Area}} = \dfrac{\text{Newton}}{\text{m}^2}$
- $1\ \text{N} = 1\ \text{kg} \cdot \text{m/s}^2$
- 1 pascal (**Pa**) = 1 N/m^2 (SI unit)
- Other commonly used units to express gas pressures:
  - Atmosphere (atm)
  - Millimeters mercury (mm Hg)
  - Torr
  - Bar
  - Pound per square inch (psi)



Chapter 11 Section 1

Units of Pressure

TABLE 11.2 Units of Pressure Commonly Used in Chemistry

Unit	Origin	Definition
standard atmosphere (atm)	Pressure at sea level	1 atm = 101,325 Pa
mmHg	Barometer measurement	1 mmHg = 133.322 Pa
torr	Name given to mmHg in honor of Torricelli, the inventor of the barometer	1 torr = 133.322 Pa
bar	Same order of magnitude as atm, but a decimal multiple of Pa	1 bar = 1×10^5 Pa

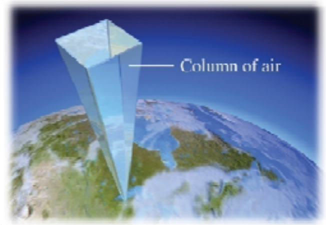
Chapter 11 Section 1

Atmospheric Pressure

- How much pressure does the air exert on Earth?

  Applying :

  $$\text{pressure} = \frac{\text{force}}{\text{area}}$$



An imaginary column of air above a 1 cm^2 spot at sea level weighs approximately 1 kg.

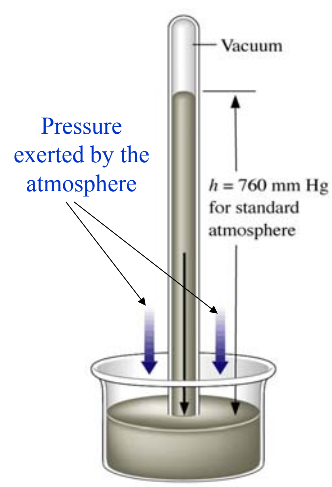
- force = 1 kg × 9.81 m/s^2 ≈ 10 kg · m/s^2 = 10 N
- Pressure = 10 N / 0.0001 m^2 = 1×10^5 Pa

1×10^5 Pa is the approximate pressure at sea level and is roughly equal to 1 atm.

Chapter 11 Section 1

Barometer

- The **barometer** is a simple device that is used to measure the atmospheric pressure and was invented by Torricelli.
- The pressure exerted by the mercury column is equal to the pressure exerted by the atmosphere.

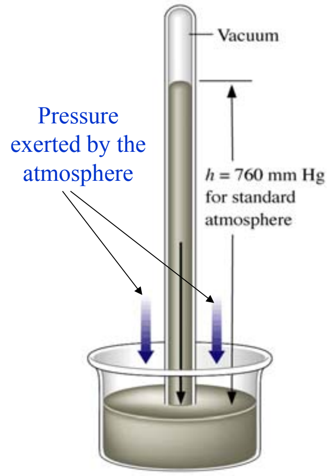


Chapter 11 Section 1

Standard Atmospheric Pressure

- ***Standard atmospheric pressure*** (1 atm) is defined as the pressure that would support a column of mercury exactly 760 mm high at 0˚C at sea level.

1 atm*
101,325 Pa
760 mmHg*
760 torr*
1.01325 bar
14.7 psi
*These are exact numbers.

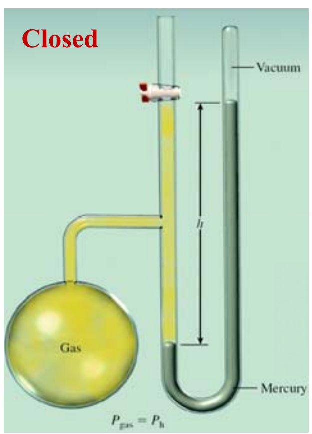


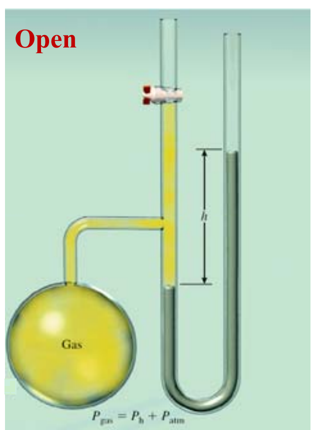
1 atm ≡ 760 mmHg

Chapter 11 Section 1

Manometer

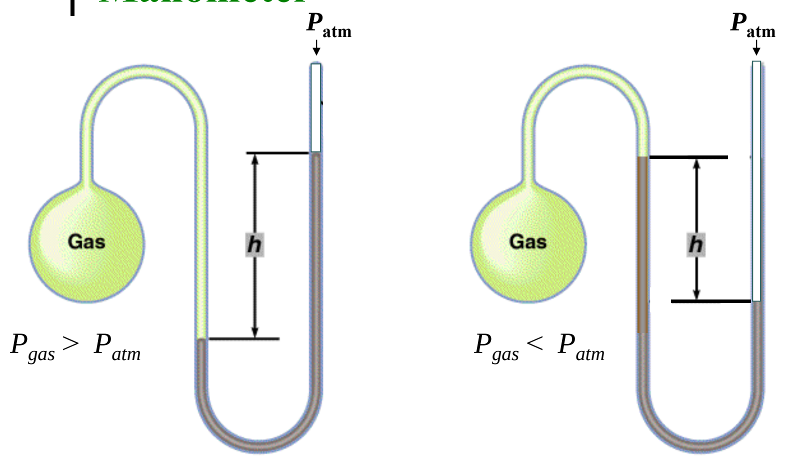
- A ***manometer*** is a device used to measure pressures other than atmospheric pressure; e.g. gas samples.





Chapter 11 Section 1

Manometer



$$P_{gas} = P_{atm} + h \qquad P_{gas} = P_{atm} - h$$

Try: http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/gaslaw/manometer4-1.html

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Chapter 11 Section 1

Exercise

If a weatherman says that atmospheric pressure is 29.12 inches of mercury, what is it in torr?

$$29.12\,\text{in}\left(\frac{2.54\,\text{cm}}{1\,\text{in}}\right)\left(\frac{10\,\text{mm}}{1\,\text{cm}}\right)\left(\frac{1\,\text{torr}}{1\,\text{mm}}\right) = 739.6\,\text{torr}$$

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The Gas Laws

- Boyle's Law

  $PV = k$ at constant T and n

- Charles's and Gay-Lussac's Law

  $V = bT$ at constant P and n

- Avogadro's Law

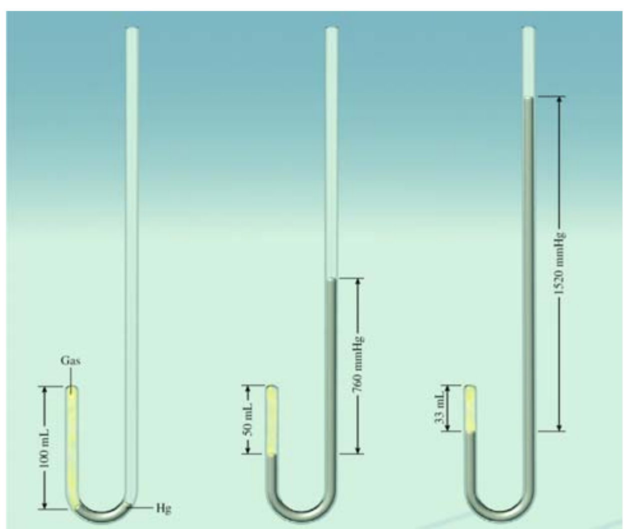
  $V = an$ at constant T and P

Where V is volume, P is pressure, T is temperature, and n is number of moles.

Boyle's Law

- Boyle studied the relationship between pressure (P) and volume (V) for gases at constant temperature (T).

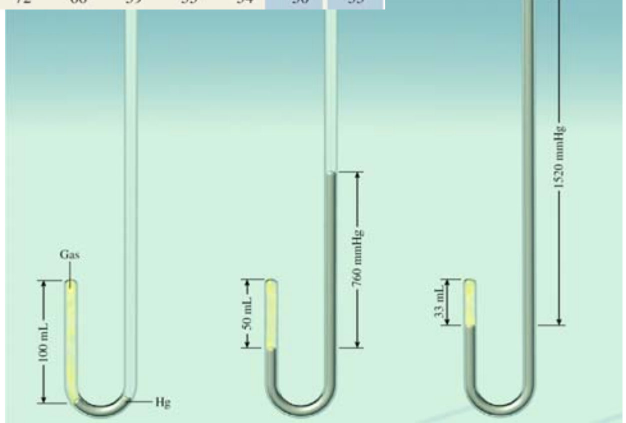
- Adding more Hg at constant T compresses the gas (less V and higher P).



Chapter 11 Section 2

Boyle's Law

TABLE 11.3	Typical Data from Experiments with the Apparatus of Figure 11.7									
P (mmHg)	760	855	950	1045	1140	1235	1330	1425	1520	2280
V (mL)	100	89	78	72	66	59	55	54	50	33



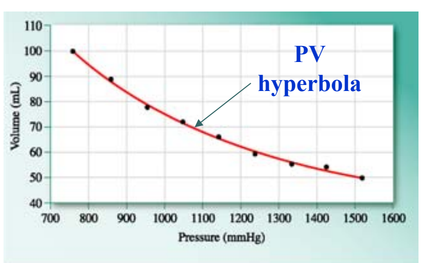
V* α 1/*P

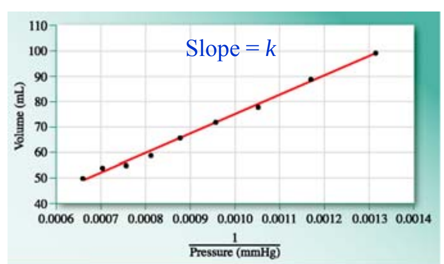
Boyle's Law

at a constant T and for a fixed quantity of gas

Chapter 11 Section 2

Plotting Boyle's Results





Inverse relationship.

$V \propto 1/P$

P drops by half when V is doubled.

$PV = \text{constant}$

$y = mx + b$

(equation of the straight line)

$$V = \frac{k_1}{P}$$

Boyle's Law

$$PV = k_1$$

- The product of P and V is always equal to the same constant as long as the temperature is held constant and the amount of gas doesn't change.

Application of Boyle's Law

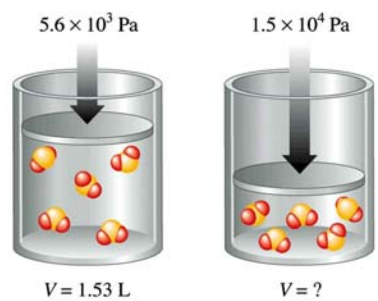
1.53 L of SO_2 is initially at a pressure of 5.6×10^3 Pa. If the pressure is changed to 1.5×10^4 Pa at constant temperature, what will be the new volume?

$PV = k_1$

$P_1V_1 = k_1 = P_2V_2$

T is constant and the gas is assumed to be ideal.

$$V_2 = \frac{P_1V_1}{P_2} = 0.57\text{ L}$$



Charles's and Gay-Lussac's Law

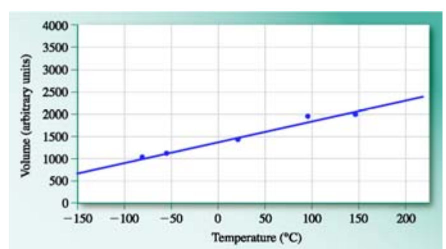
Pouring liquid nitrogen (-196˚C) lowers the temperature of the He balloon; which causes the volume to decrease.

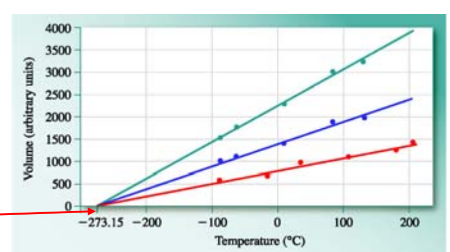
Charles's and Gay-Lussac's Law

- Charles and Gay-Lussac found that the *V* of a gas increases linearly with the increase of *T*, when *P* is held constant.

$$\boxed{V = k_2 T}$$

- The volume-temperature lines having different slopes at different pressure all extrapolate to zero volume at T = -273.15 ˚C.



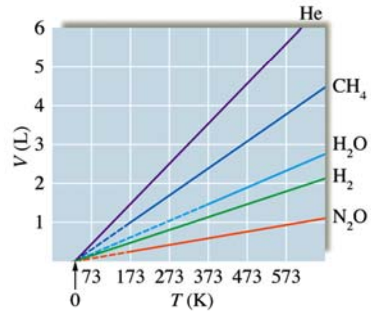


Chapter 11 Section 2

Absolute Zero

- $T = -273.15$ °C was defined by Lord Kelvin to be the ***absolute zero***.
- Absolute zero is common for all gases.
- Absolute temperature scale:

  K = °C + 273.15
- 1×10^{-6} K was reached in laboratories, but 0 K has never been reached.



Absolute zero

Chapter 11 Section 2

Charles's Law

$$V = k_2 T$$

- The volume of a fixed amount of gas maintained at a constant pressure is directly proportional to the absolute temperature of the gas.

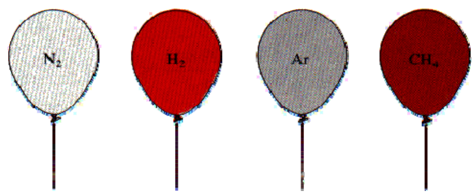
$$\frac{V_1}{T_1} = k_2 = \frac{V_2}{T_2}$$

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

Chapter 11 Section 2

Avogadro's Law

- ***Avogadro's law***: Equal volumes of different gases contain the same number of particles at the same T and P.



$$\boxed{V = k_3\, n}$$

(Linear relationship between V and n at constant P and T)

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

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Chapter 11 Section 2

Applications of Avogadro's Law

- A balanced chemical equation reveals the ratio of reactants and products in terms of *number of moles* as well as in terms of *volumes* at constant T and P.

$$3H_2(g) \quad + \quad N_2(g) \quad \longrightarrow \quad 2NH_3(g)$$

3 molecules	+	1 molecule	$\longrightarrow$	2 molecules
3 moles	+	1 mole	$\longrightarrow$	2 moles
3 volumes	+	1 volume	$\longrightarrow$	2 volumes

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Applications of Avogadro's Law

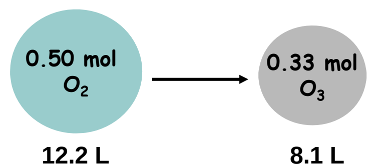
12.2 L sample containing 0.50 moles of O_2 at $P = 1.00$ atm and $T = 25^{\circ}C$ is converted into O_3 at the same T and P. What would be the volume of O_3?

$$3O_2 \longrightarrow 2O_3$$

$$\text{\# mol of } O_3 \text{ produced} = 0.50 \cancel{\text{mol } O_2} \times \frac{2 \text{ mol } O_3}{3 \cancel{\text{mol } O_2}} = 0.33 \text{ mol } O_3$$

$$\frac{n_{O_2}}{V_{O_2}} = k_3 = \frac{n_{O_3}}{V_{O_3}}$$

$$V_2 = \left(\frac{n_2}{n_1}\right) V_1 = 8.1 \text{ L}$$



The volume decreases because fewer number of molecules will be present after O_2 is converted to O_3.

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Chapter 11 Section 2

Exercise

How will the volume of a given gas change if the quantity of gas, absolute temperature, and pressure, all double?

Avogadro	Charles	Boyle
× 2	× 2	× 1/2

$$2 \times 2 \times \frac{1}{2} = 2 \quad \Rightarrow \text{ volume doubles}$$

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Chapter 11 Section 3

The Ideal Gas Equation

		Constant
Boyle's law:	$V = k_1 \, 1/P$	T and n
Charles's Law:	$V = k_2 \, T$	P and n
Avogadro's Law:	$V = k_3 \, n$	T and P

$$V \propto \frac{n\,T}{P} \quad \Rightarrow \quad V = R\frac{n\,T}{P}$$

R is the ***gas constant***

$$R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}$$

Ideal Gas Equation $\boxed{PV = nRT}$

ET D'HISTOIRE NATURELLE. 131

MÉMOIRE

SUR LES MASSES RELATIVES

DES MOLÉCULES DES CORPS SIMPLES,

Ou densités présumées de leurs gaz, et sur la constitution de quelques-uns de leurs composés, pour servir de suite à l'*Essai* sur le même sujet, publié dans le *Journal de Physique*, juillet 1811;

PAR A. AVOGADRO.

Communiqué à M. Delamétherie, en janvier 1814.

Dans mon *Essai d'une manière de déterminer les Masses relatives des Molécules des corps*, etc., j'ai proposé une hypothèse très-naturelle, ce me semble, et à laquelle on n'a rien substitué jusqu'ici, pour expliquer le fait découvert par M. Gay-Lussac, que les volumes des substances gazeuses qui se combinent entre elles, et des gaz composés qui en résultent, sont toujours dans des rapports très-simples entre eux. Cette hypothèse consiste à dire, que les volumes égaux de substances gazeuses, à pression et température égales, représentent des nombres égaux de molécules, ensorte que les densités des différens gaz sont la mesure des masses des molécules propres à ces gaz, et que les rapports des volumes dans les combinaisons ne sont que les rapports entre les nombres de molécules qui se combinent pour former des molécules composées.

Cette hypothèse, une fois admise, en confirmant en partie les résultats auxquels MM. Dalton, Davy et autres ont été conduits par des considérations particulières sur les masses des molécules des différentes substances connues, d'après leurs pro-

B 2

Chapter 11 Section 3

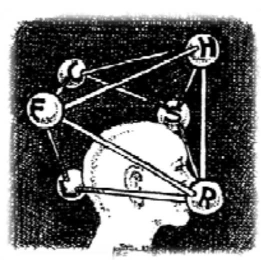
The Ideal Gas Equation

- It is an equation of state for a gas. A particular state of a gas is described by its P, V, n and T.
- A gas that precisely obeys the ideal gas law is said to be "ideal", or to behave "ideally". An ***ideal gas*** is a hypothetical substance.
- We always assume ideal gas behavior when you solve problems involving gases.
- The gas constant, R, can be expressed in several ways.

Numerical Value	Unit
0.08206	L · atm/K · mol
62.36	L · torr/K · mol
0.08314	L · bar/K · mol
8.314	m^3 · Pa/K · mol
8.314	J/K · mol
1.987	cal/K · mol

$$\boxed{P \cdot V = n \cdot R \cdot T}$$

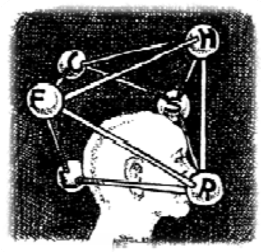
P = pressure
V = volume
n = number of molecules
R = constant
T = temperature



Chapter 11 Section 3

Applications of the Ideal Gas Equation

$\boxed{P \cdot V = n \cdot R \cdot T}$
P = pressure
V = volume
n = number of molecules
R = constant
T = temperature



For an ideal gas, calculate the pressure of the gas if 0.215 mol occupies 338 mL at 32.00ºC.

☞ Don't forget to use proper units.

n = 0.215 mol

V = 338 mL = 0.338 L

T = 32.00 + 273.15 = 305.15 K

P = ?

$PV = nRT$ => $P = nRT/V$

$$P = \frac{(0.215\ \text{mol})\left(0.08206\ \frac{\text{L}\times\text{atm}}{\text{mol}\times\text{K}}\right)(305.15\ \text{K})}{0.338\ \text{L}} = 15.928$$

$= 15.9$ atm

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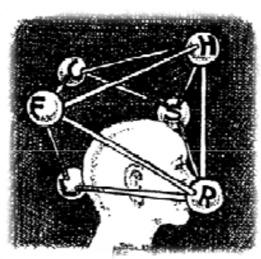
Chapter 11 Section 3

Applications of the Ideal Gas Equation

$\boxed{P \cdot V = n \cdot R \cdot T}$
P = pressure
V = volume
n = number of molecules
R = constant
T = temperature



A sample of H_2 gas has $V = 8.56$ L @ $T = 0.00$ ºC and $P = 1.5$ atm. How many H_2 molecules are present?

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Applications of the Ideal Gas Equation

A sample methane with $V = 3.8$ L @ 5ºC is heated to 86ºC at constant P, what is the new volume?

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

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$ (Charles's Law)

$V_2 = V_1T_2 / T_1$ = (359 $\cancel{\text{K}}$)(3.8 L) / 278 $\cancel{\text{K}}$ = 4.9 L

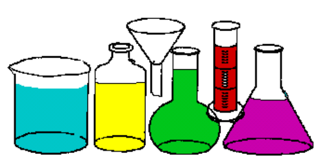
- Again, don't forget to state T in K.
- Does the answer make sense to you?

Standard Temperature and Pressure

- The condition of a sample of gas with:
  - $T = 0$˚C, and
  - $P = 1$ atm

  is know as ***standard temperature and pressure (STP)***.

- The ideal gas equation is *not* exact, but for most of the real gases, it is quite accurate near STP.



Chapter 11 Section 3

Molar Volume

What is the volume of 1.000 mole of a gas at standard temperature and pressure (STP)?

STP : $T = 0°C$; $P = 1$ atm.

$$V = \frac{nRT}{T} = \frac{(1.000 \text{ mol})(0.0821 \text{ L} \cdot \text{atm / K} \cdot \text{mol})(273.15 \text{ K})}{(1.000 \text{ atm})}$$

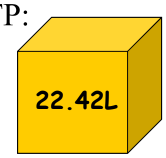
$= 22.42$ L (Molar volume @ STP)

TABLE 5.2 Molar Volumes for Various Gases at 0°C and 1 atm

Gas	Molar Volume (L)
Oxygen (O_2)	22.397
Nitrogen (N_2)	22.402
Hydrogen (H_2)	22.433
Helium (He)	22.434
Argon (Ar)	22.397
Carbon dioxide (CO_2)	22.260
Ammonia (NH_3)	22.079

☞ For an ideal gas @ STP:

1 mol ≡ 22.42 L



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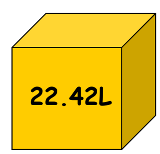
Chapter 11 Section 3

Applications of Molar Volume

At STP, how many atoms of neon gas are present in 0.500 L sample of neon gas?

Assuming ideal behavior:

1 mol Ne has volume of 22.42 L



$$\text{\# mol Ne} = 0.500 \text{ L Ne} \times \frac{1 \text{ mol Ne}}{22.42 \text{ L Ne}}$$

$$= 2.23\times10^{-2} \text{ mol Ne}$$

$$\text{\# Ne atoms} = 2.23\times10^{-2} \text{ mol Ne} \times \frac{6.022\times10^{23} \text{ Ne atoms}}{1 \text{ mol Ne atoms}}$$

$$= 1.34\times10^{22} \text{ Ne atoms}$$

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Chapter 11 Section 3

Molar Mass of a Gas

For an ideal gas:

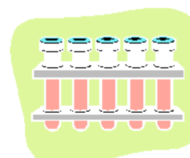
$$n = \frac{\text{mass}}{MM} \quad \text{where } MM \text{ is the molar mass.}$$

$$P = \frac{nRT}{V} = \frac{(\text{mass}/MM)RT}{V} = \frac{(\text{mass})RT}{V\,(MM)}$$

Since the density of a gas (d) = $\frac{\text{mass}}{V}$

Then, $P = d\,\frac{RT}{MM}$ or $\boxed{MM = d\,\frac{RT}{P}}$

g/L



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Chapter 11 Section 4

Reactions with Gaseous Reactants and Products

$$\underset{152\text{ g}}{CaCO_3\,(s)} \xrightarrow[\text{Heat}]{@STP} CaO\,(s) + \underset{V=?}{CO_2\,(g)}$$

g of $CaCO_3$ → mol of $CaCO_3$ → mol of CO_2 → Molar Volume

Calcite/limestone

$CaCO_3$

Quicklime (CaO)

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Chapter 11 Section 4

Reactions with Gaseous Reactants and Products

$$\underset{?\,g}{2NaN_3\,(s)} \xrightarrow{@\ 25^{o}C \text{ and } 1.0 \text{ atm}} 2Na\,(s) + \underset{V = 65L}{3N_2\,(g)}$$

NaN_3 (sodium azide) is an explosive material. It explodes very fast and completes the reaction in 40 ms. Thus, it is used in airbag technology.

Chapter 11 Section 4

Reactions with Gaseous Reactants and Products

$$2LiOH\,(aq) + CO_2(g) \xrightarrow{@\ 25^{o}C \text{ and } 2.5\times10^{5}\,L} Li_2CO_3(s) + H_2O\,(l)$$

The pressure of the submarine drops from 0.9970 atm to 0.9891 atm as a result of CO_2 being consumed by LiOH scrubber. How many grams of CO_2 are consumed?

At constant temperature and volume:

$$n = P \times \left(\frac{V}{RT}\right)$$

$$\Delta n = \Delta P \times \left(\frac{V}{RT}\right)$$

$$\Delta n_{CO_2} = 7.9\times10^{-3}\,\text{atm}\,\frac{2.5\times10^{5}\,\text{L}}{\left(0.08206\,\frac{\text{L}\cdot\text{atm}}{\text{K}\cdot\text{mol}}\right)(298.15\text{K})} = 81\,\text{mol}\,CO_2$$

Mass of CO_2 = (81 mol) × (44.01g/mol) = 3.6×10^{3} g CO_2

Reactions with Gaseous Reactants and Products

	$CH_4\ (g)$ +	$2O_2\ (g)$ $\longrightarrow$	$CO_2\ (g)$ + $H_2O\ (g)$
Pressure	1.65 atm	1.25 atm	2.50 atm
Volume	2.80 L	35.0 L	?
Temperature	298 K	304 K	398 K
$n = \dfrac{PV}{RT}$	0.189 mol	1.75 mol =>	0.189 mol

Limiting reactant because it requires
0.189 mol × 2 = 0.378 mol of O_2

$$V_{CO_2} = \frac{nRT}{P} = \frac{(0.189\text{ mol})(0.0821\text{ L}\cdot\text{atm/K}\cdot\text{mol})(398\text{K})}{(2.50\text{ atm})} = 2.47\text{ L } CO_2$$