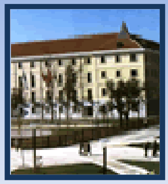


TOPIC 2

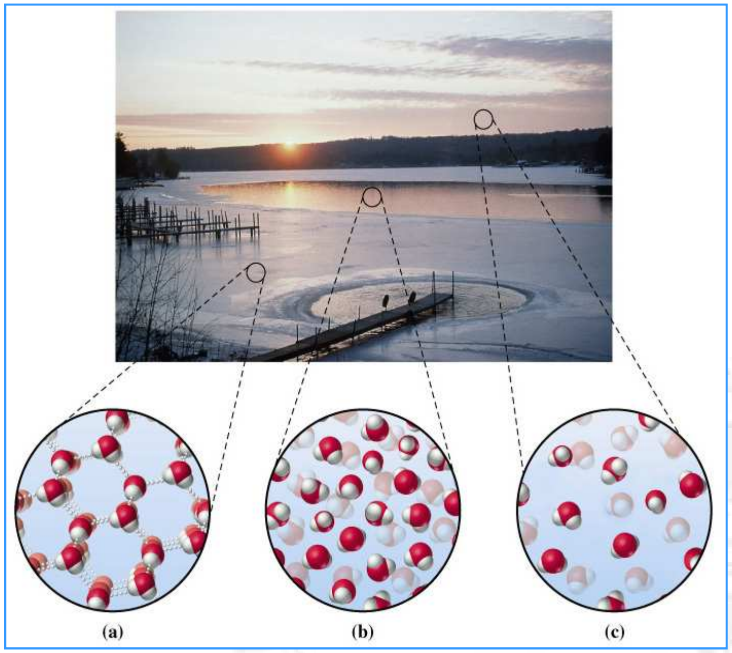
States of Matter (I) - Gases

Contents

1. Introduction
2. Pressure measurement
3. The Ideal Gas equation
4. Efusion and Diffusion
5. Kinetic Molecular Theory
6. Dalton’s Law of Partial Pressure
7. Real Gases: van der Waals equation

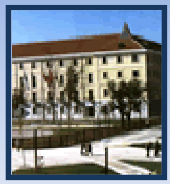
1. Introduction

States of Matter



General Chemistry: Principles and Moderns Applications
R.H. Petrucci

	Distances	Interactions	Movement
Solid	Short	Very strong	Highly restricted
Liquids	Short	Strong	Restricted
Gases	Long	Weak	Almost Free



1. Introduction

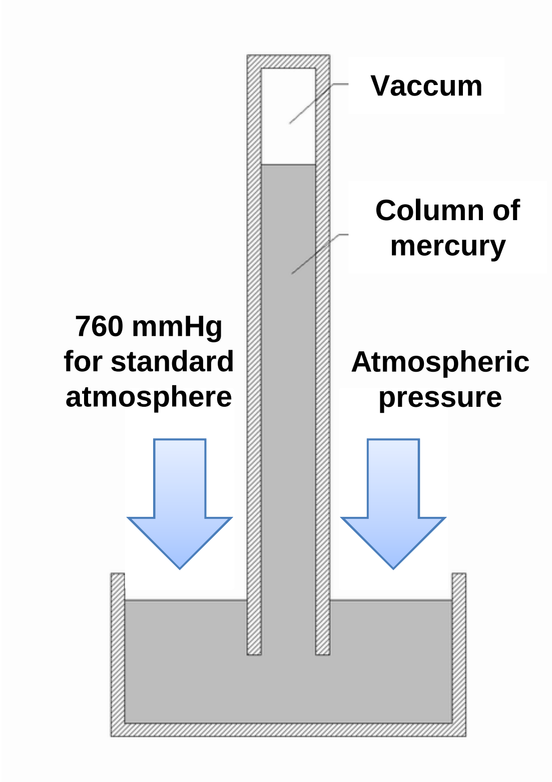
Where do we find gases?

http://en.wikipedia.org/wiki/File:Michael_Schumacher_Ferrari_2004.jpg

http://en.wikipedia.org/wiki/File:Automobile_exhaust_gas.jpg

2. Pressure Measurement

Barometer



P = d·g·h

d = density
g = gravity
h = height

1 atm = 760 mmHg ⇒ atmospheric pressure at sea level

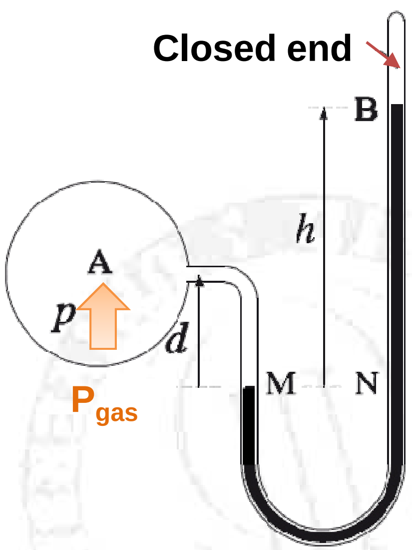
1 Torr = 1 mmHg

Pa = N/m^2 <> 1/133.322 Torr

Standard Atmospheric Pressure
1.00 atm
760 mm Hg, 760 Torr
101.325 kPa
1.01325 bar
1013.25 mbar

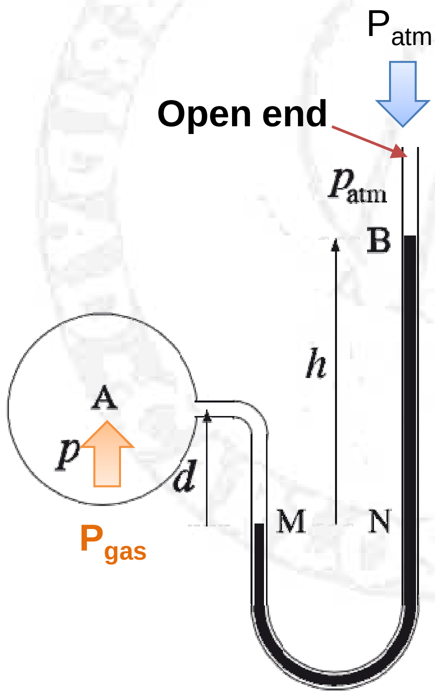
2. Pressure Measurement

Manometer



(a)

$P_{gas} = P_h$



(b)

$P_{gas} = P_{atm} + P_h$

$P_h = \rho_{(Hg)} \cdot g \cdot h$

$\rho_{(Hg)}$ = density of mercury 13.6 g·cm^3.

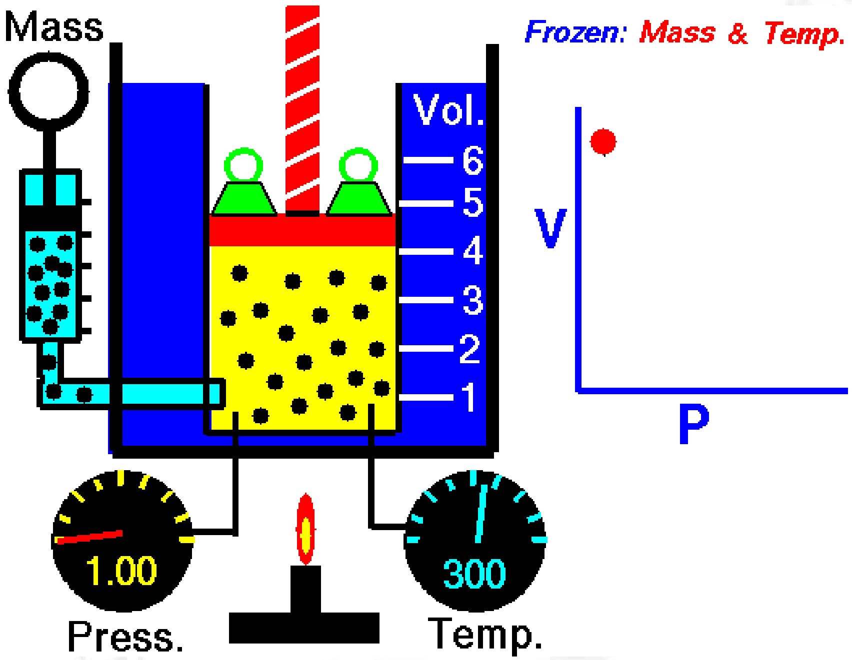
g = 9.8 m·s^{-2}.

h = distance B-N (in m).

3. Gases Laws

Boyle's Law

$P \propto 1/V$ when T & n are constant

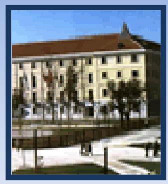


http://www.grc.nasa.gov/WWW/K-12/airplane/boyle.html

$P \cdot V = k$

$P_1 \cdot V_1 = P_2 \cdot V_2 = P_3 \cdot V_3 = \cdots$

For a fixed amount of an ideal gas at a constant temperature, pressure and volume are inversely proportional.



3. Gases Laws

Charles and Gay-Lussac’s Law

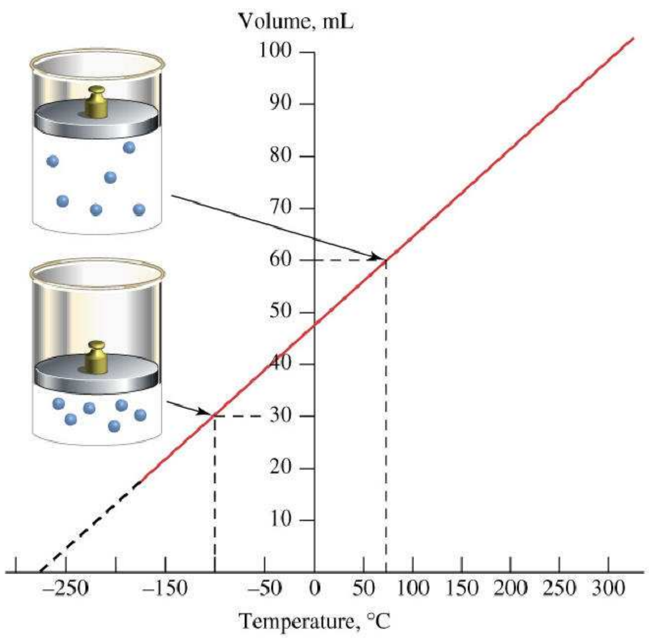


Image from:
http://www.unizar.es/lfnae/luzon/CDR3/charles.jpg

$V \propto T$

$V_1/T_1=V_2/T_2=k$ **(P & n constant)**

$V = a(t + 273)$ where:

- **V=volume**
- **t=temperature in Celsius**
- **a= slope of the straight line.**

$T = 273.15 + t$

T= Absolute temperature (K)

For a fixed amount of an ideal gas at a constant pressure, the volume is directly proportional to the temperature. **$V = c \cdot T$**

3. Gases Laws

The Ideal Gas equation

- Boyle's law $V \propto 1/P$
- Charles's law $V \propto T$
- Avogadro's law $V \propto n$

$$V \propto \frac{nT}{P}$$

PV = nRT

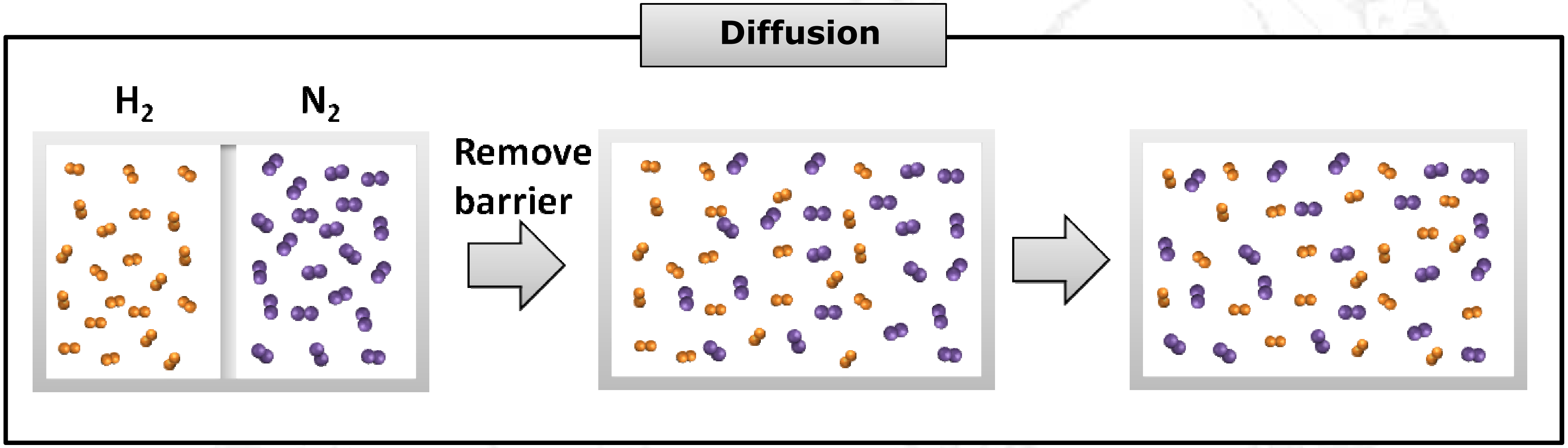
The Gas Constant

At STP (0 ºC and 1atm) 1 mol gas = 22.4 L gas

$$R = \frac{pV}{nT} = \frac{1atm * 22.4L}{1mol * 273.15K} = 0.082\frac{atmL}{Kmol}$$

$R = 8{,}206 \cdot 10^{-2}$ atm·l·mol^{-1}·K^{-1}
<> 8.314 J·mol^{-1}·K^{-1} <> 1,987 cal·mol^{-1}·K^{-1}

4. Diffusion and Efusion



Graham's Law

Graham's law states that the rate of diffusion of a certain gas is inversely proportional to the square root of its molecular weight.

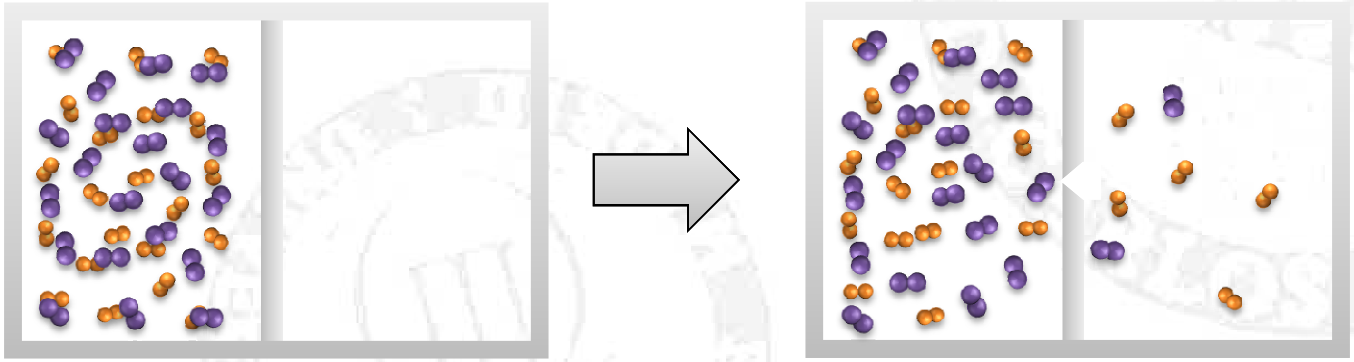
$$R \propto \sqrt{\frac{1}{\rho}}$$

$$R \propto \sqrt{\frac{1}{M}}$$

4. Diffusion and Efusion

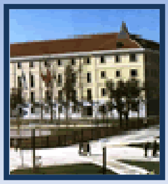
Why do tyres deflate?

Efusion



$$\frac{\text{rate of effusion of A}}{\text{rate of effusion of B}} = \frac{R_A}{R_B} = \sqrt{\frac{M_B}{M_A}}$$

Graham's Law



5. Kinetic Molecular Theory

Why do tyres deflate?

- Particles are point masses (with no volume) in constant, random and straight line motion.
- Particles are separated by great distances, almost infinite.
- Collisions are elastic.
- There are no intermolecular forces between the particles.
- The total energy of the system remains constant.

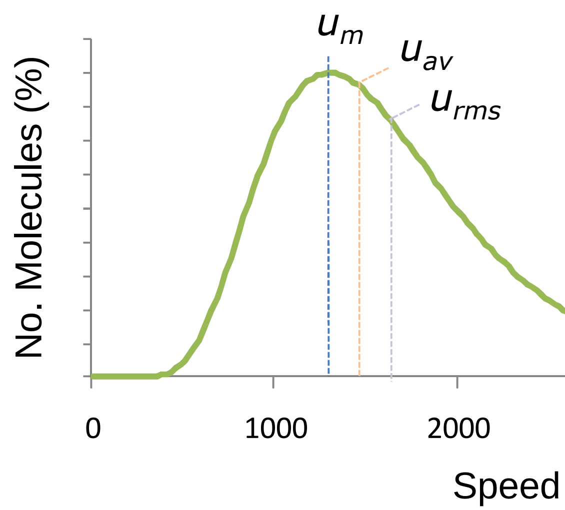
$$\text{Probability} \propto \frac{N}{V}$$

$$p = \frac{Nm\overline{u^2}}{3V}$$

$$\text{Concentration} = \frac{N}{V}$$

$$\text{Energy per collision} = m\overline{u^2}$$

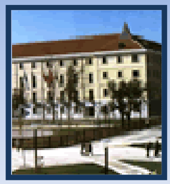
$$\text{Degrees of liberty} = 3$$



u_m is the modal speed

u_{av} is the simple average

$u_{rms} = \sqrt{\overline{u}^2}$



5. Kinetic Molecular Theory

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$

Let's consider 1 mol of gas:

$$PV = \frac{1}{3} N_A m\overline{u}^2$$

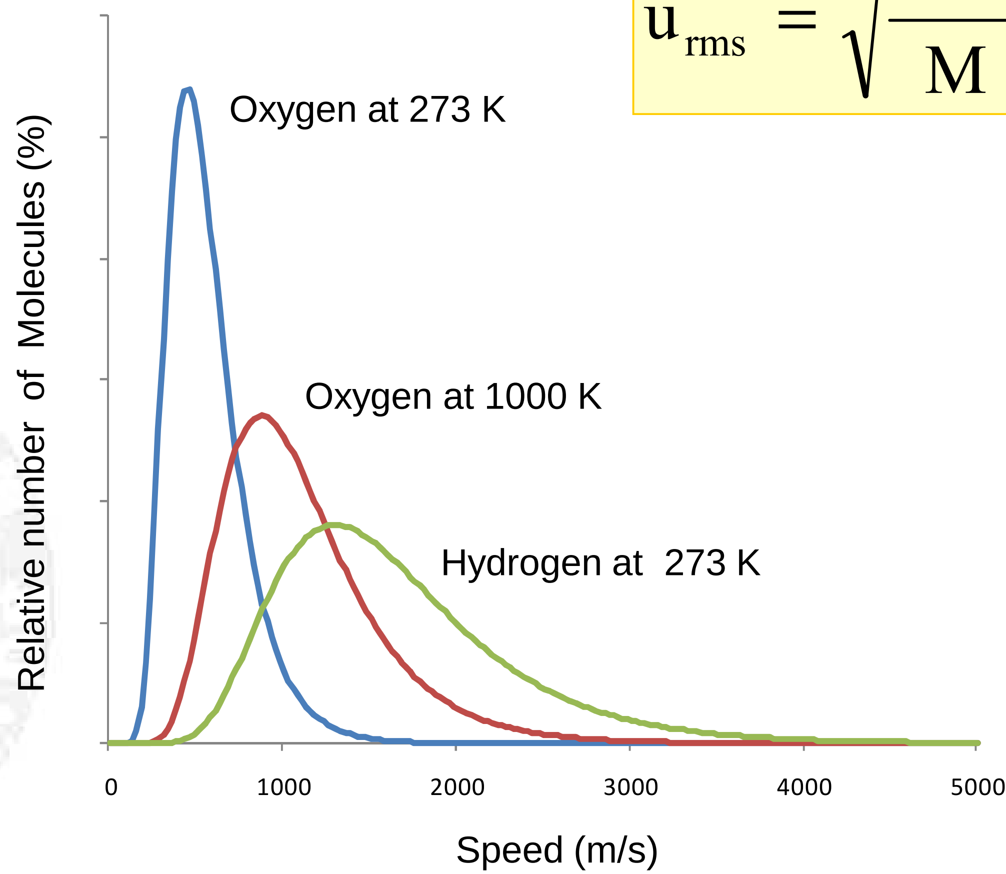
If **PV = RT** (for n= 1 mol):

$$3RT = N_A m\overline{u}^2$$

Considering that $\boldsymbol{N_A m = M}$:

$$3RT = M\overline{u}^2$$

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$



6. Dalton's Law of Partial Pressure

The total pressure of a mixture of gases is the sum of the pressures that each gas would exert if it were present alone.

$$p_T = p_A + p_B + p_C + \cdots = \sum_{i=1}^{N} p_i$$

$$p_i = \frac{n_i RT}{V}$$

$$p_i = \chi_i p_T$$ V , T constant

$where:$

$$\chi_i = \textbf{mole fraction}$$

$$p_T = \textbf{total pressure}$$

Example: Calculate the total pressure of a mixture of H_2 and O_2 whose partial pressures are P_{H2} = 2.9 atm and P_{O2} = 7.2 atm respectively

P_1 = 3.1 atm

P_2 = 6.4 atm

$P_{total} = P_1 + P_2 = 9.5$ atm

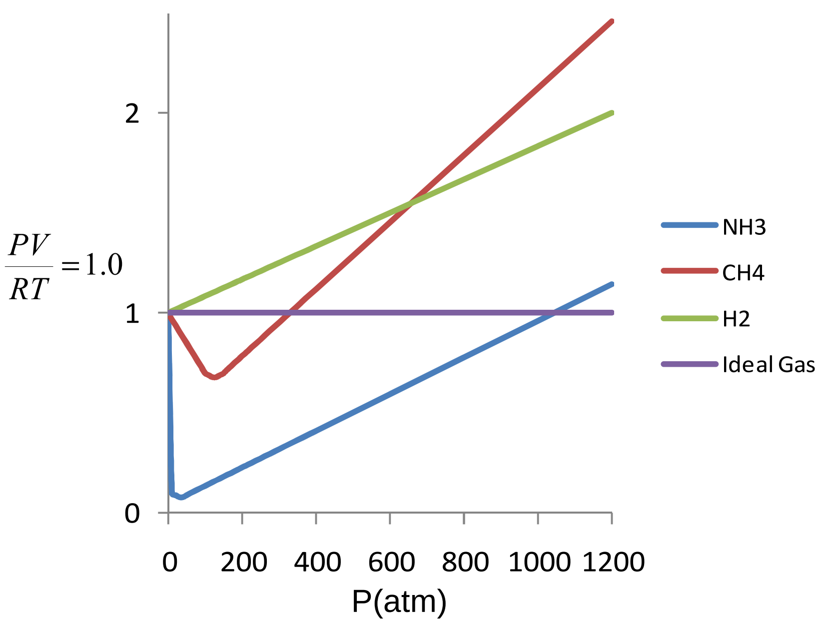
7. Real Gases

For exactly one mole of an ideal gas: $$\frac{PV}{RT} = 1.0$$

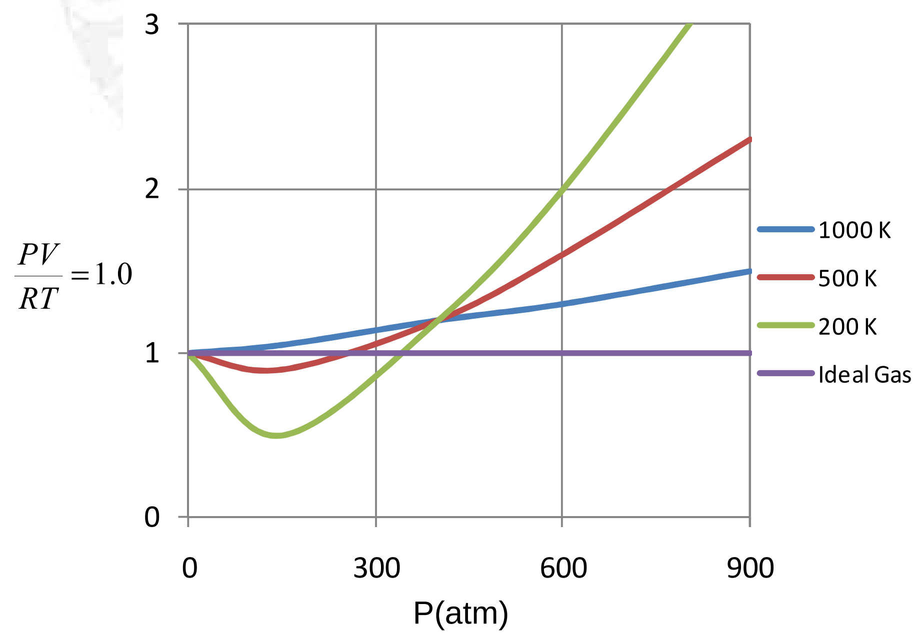
The factor "Z = PV/NRT" is known as **COMPRESSIBILITY FACTOR** and should be one for an ideal gas.

For exactly one mole of a real gas:

a) Effect of the nature of the gas



b) Effect of the temperature



7. Real Gases

$$P \cdot V = n \cdot R \cdot T$$

The Ideal Gas Equation

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

The van der Waals Equation
(a real gas equation)

Pressure Correction

Volume Correction

a: Effect of intermolecular forces
(i.e. nature of the gas)

b: Effect of molecular volume → Particles are NOT point masses (with no volume as states the kinetic molecular theory).