Ideal Gas Law(s)

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What is pressure?

A simple definition of pressure (P) is the collision of gas particles with the walls of the container.

If we say that each collision strikes the wall with a certain force (F) over a particular area of the wall (A):

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

There are many units for pressure, but you should be comfortable with four of these:

Unit		Value
Atmosphere	(atm)	1 atm
Millimeter of mercury	(mm Hg)	1 atm = 760 mm Hg
Torr		1 atm = 760 Torr
Bar		1 atm = 1.01325 bar

GASES

1. Gases take up the volume of the container – has no definite shape or volume
2. Gases mix well – diffusion
3. Gases exert pressure

THINGS WE CARE ABOUT FOR GASES

- Pressure (P)
- Volume (V)
- Temperature (T)
- Moles (n)

We'll come back to these in a moment.

ATMOSPHERIC PRESSURE

Remember that we are always under the pressure of the atmosphere, which is defined as 1 atm.

Any system that is allowed to equilibrate with the pressure of the atmosphere will try to obtain atmospheric pressure.

This is how balloons work because they can change their volume to maintain atmospheric pressure *inside*.

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- When the temperature is extremely low because the gas particles need some amount of kinetic energy to move around and exert pressure.

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Ideal conditions result in gas particles hitting each other less often so we can get around having to deal with the *intermolecular, attractive* forces between the particles.

The Ideal Gas Law

$$PV = nRT$$

P = absolute pressure (units: atm)

V = volume (units: L)

n = number of moles (units: mol)

T = absolute temperature (units: K)

R = universal gas constant $\left(0.08206\,\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)$

REFERENCE POINTS FOR GASES

Standard Temperature and Pressure (STP) : P = 1 atm and 273 K (0 °C)

Molar Volume: volume occupied by one mole any ideal gas at STP = 22.4 L

ALWAYS WORK IN ABSOLUTE TEMPERATURE SCALE (K)!

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n = 4.00 mol

T = 22 °C = 295.15 K

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

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$$V = 96.9\ \text{L}$$

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$$500.0\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}} \times \frac{1\ \text{mol}}{22.4\ \text{L}} = 0.02232\ \text{mol gas}$$

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Alternatively, use the Ideal Gas Law:

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

$$= \frac{(1.0\ \text{atm})\left(500.0\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}}\right)}{\left(0.08206\ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(273.15\ \text{K})}$$

$$n = 0.02232\ \text{mol gas}$$

Deriving the P vs. n relationship

Start by understanding that for us to relate P and n, T and V must both be constant.

So, we can organize the initial (P_1, V_1, n_1, V_1) and final (P_2, V_2, n_2, V_2) conditions of our gas system:

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Deriving the all the other gas laws:

Work your way through these to make sure you understand why they work (like I did on the previous slide)

Volume vs. Moles (V vs. n)

$V \propto n$ (constant T, P)

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

I		II
P_1	=	P_2
V_1		V_2
n_1		n_2
T_1	=	T_2

Change		*Constant*
V/n	=	RT/P

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

Pressure vs. Moles (P vs. n)

$P \propto n$ (constant T, V)

$$\frac{P_1}{n_1} = \frac{P_2}{n_2}$$

I		II
P_1		P_2
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Change		*Constant*
P/n	=	RT/V

$$\frac{P_1}{n_1} = \frac{RT}{V} = \frac{P_2}{n_2}$$

Pressure vs. Volume (P vs. V)

$P \propto \frac{1}{V}$ (constant n, T)

$$P_1V_1 = P_2V_2$$

I		II
P_1		P_2
V_1		V_2
n_1	=	n_2
T_1	=	T_2

Change		*Constant*
PV	=	nRT

$$P_1V_1 = nRT = P_2V_2$$

Volume vs. Temp (V vs. T)

$V \propto T$ (constant n, P)

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

I		II
P_1	=	P_2
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Change		*Constant*
V/T	=	nR/P

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P vs. V. vs. T

$PV \propto T$ (constant n)

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

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Change		*Constant*
PV/T	=	nR

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$T_1 = -25$ °C $= 284._{15}$ K $T_2 = 72$ °C $= 345._{15}$ K

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T_1 = –25 °C = 284.$_{15}$ K — T_2 = 72 °C = 345.$_{15}$ K

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$$V_2 = 8.3\ \text{L}$$

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V_2 = 70.0 L

T_2 = 50.0 °C = 323.1$_5$ K

Only n is constant.

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$$P_2 = 1.08\text{ atm}$$