

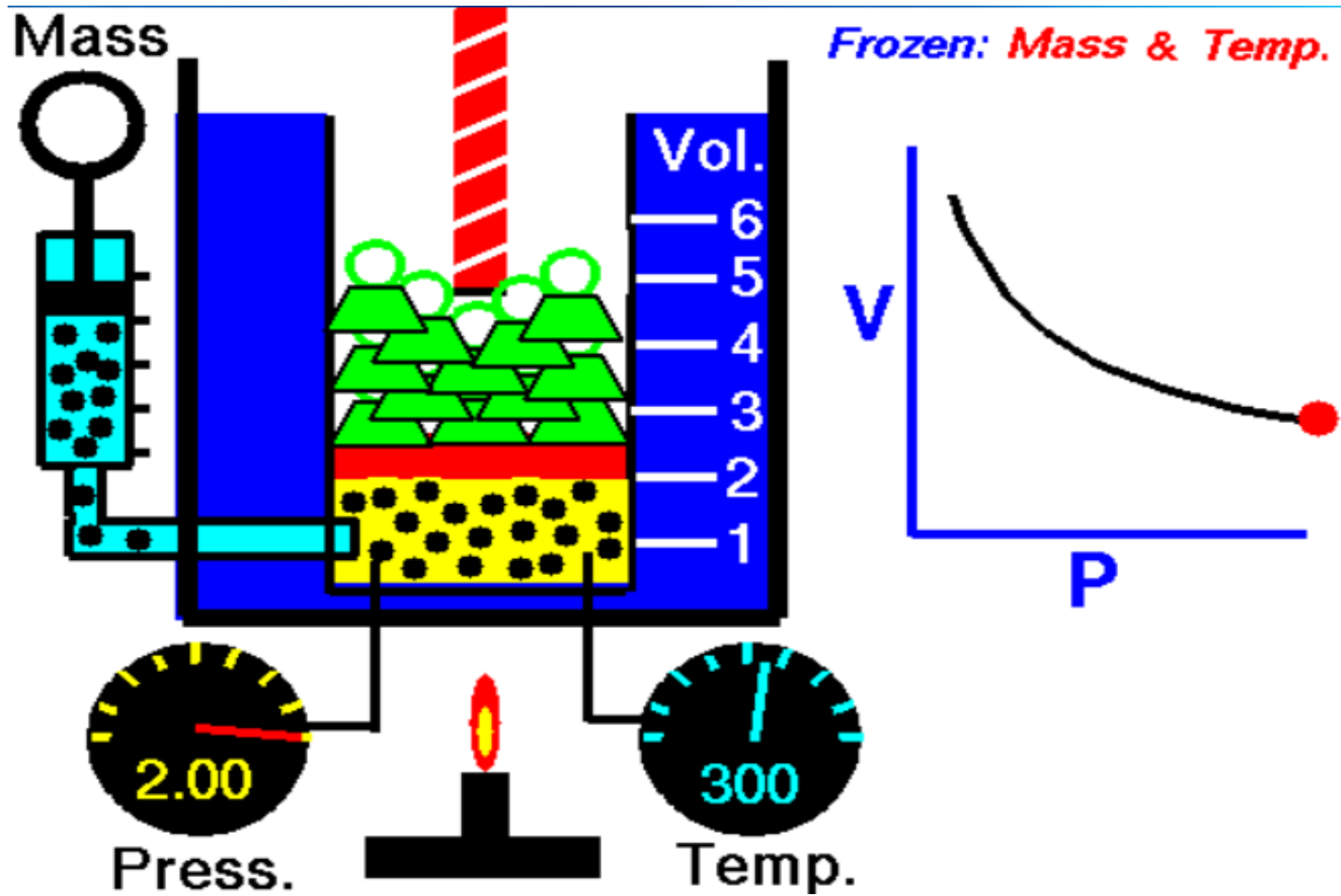
# Properties of Reservoir Fluids (PGE 362)

## Gasses

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# Gas Laws



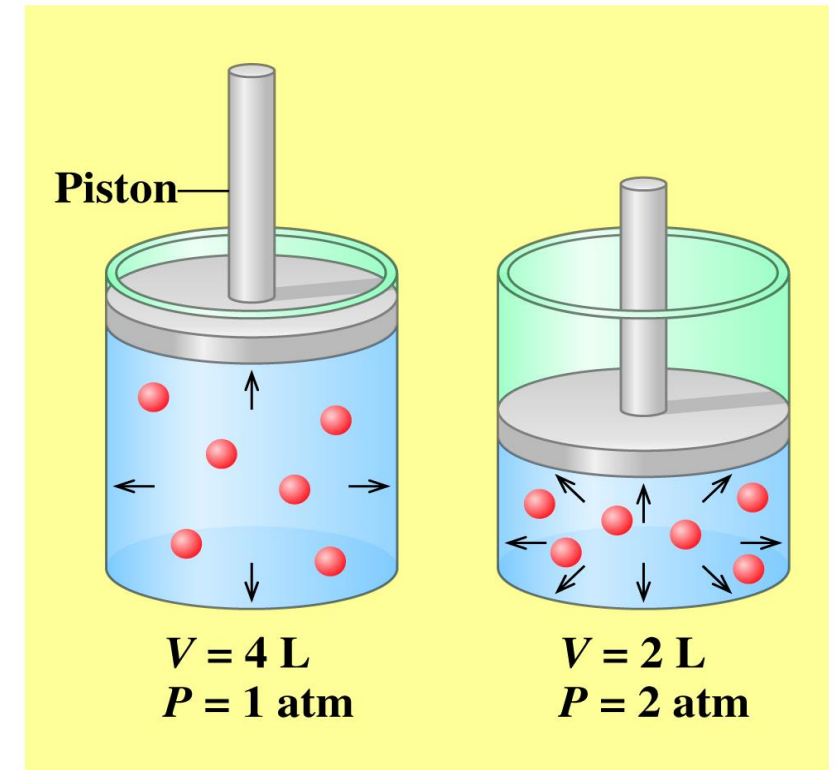
# Boyle's Law

The product of pressure and volume for a gas is constant for a fixed amount of gas at fixed temperature.

$$P \propto \frac{1}{V}$$

$$PV = \text{constant}$$

$$P_i V_i = P_f V_f \text{ (T, n are constant)}$$



# Boyle's Law

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In Boyle's Law, the product of  $PV$  is constant as long as  $T$  and  $n$  do not change.

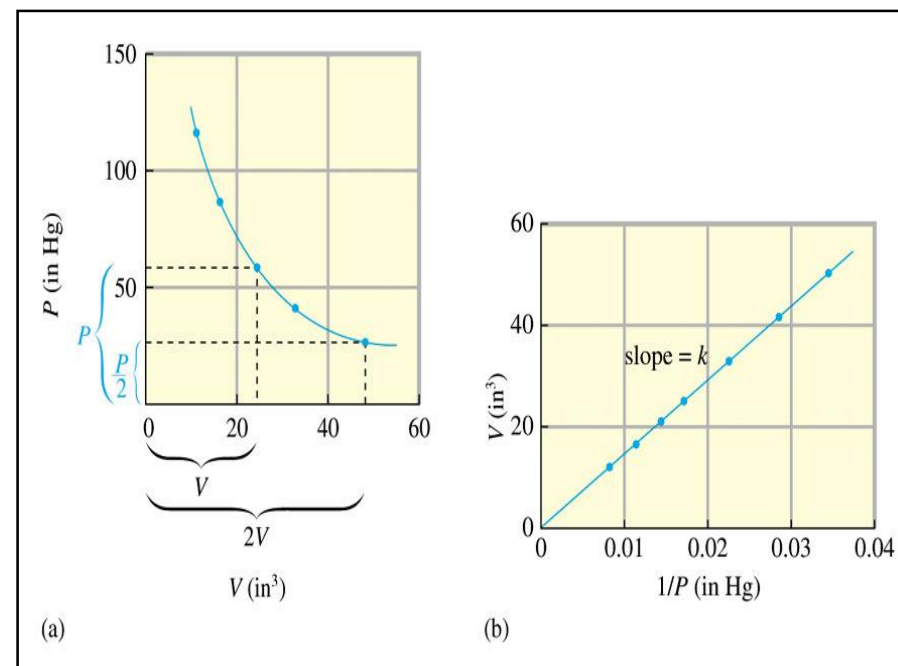
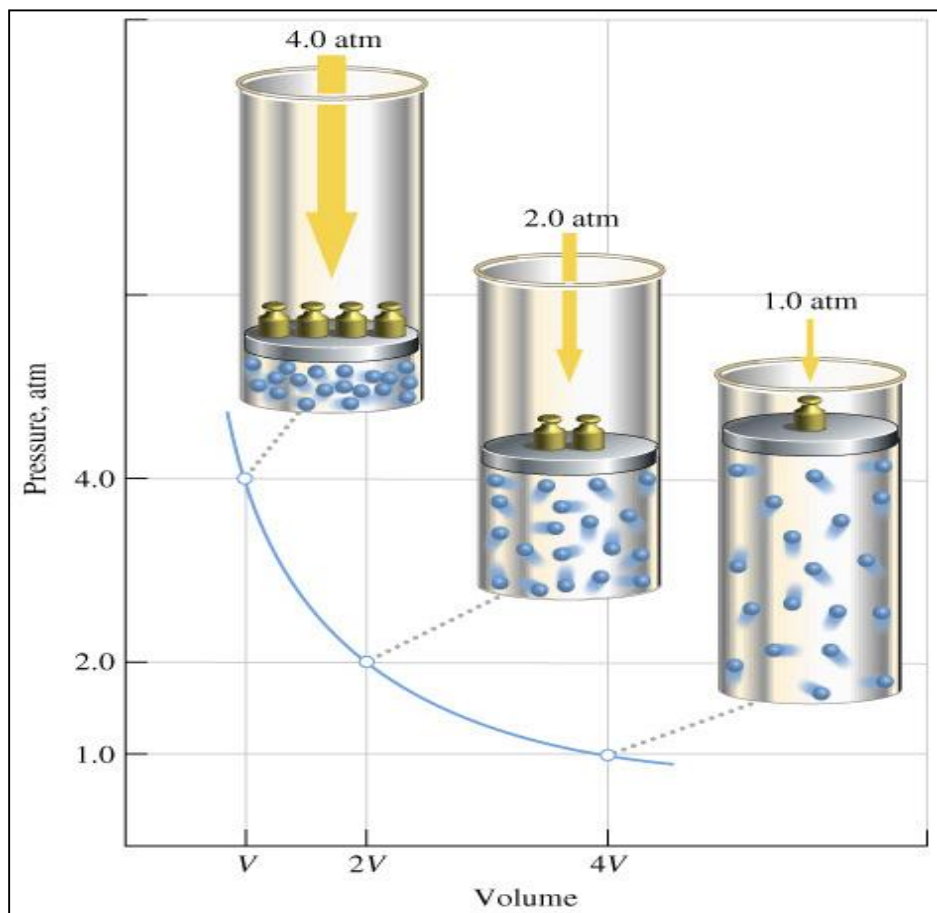
## Example:

- $P_1V_1 = 10.0 \text{ atm} \times 4.0 \text{ L} = 40.0 \text{ atm L}$
- $P_2V_2 = 8.0 \text{ atm} \times 5.0 \text{ L} = 40.0 \text{ atm L}$
- $P_3V_3 = 4.0 \text{ atm} \times 10.0 \text{ L} = 40.0 \text{ atm L}$
- $P_4V_4 = 2.0 \text{ atm} \times ?? \text{ L} = 40.0 \text{ atm L}$

## Answer :

$$V_4 = 20 \text{ L}$$

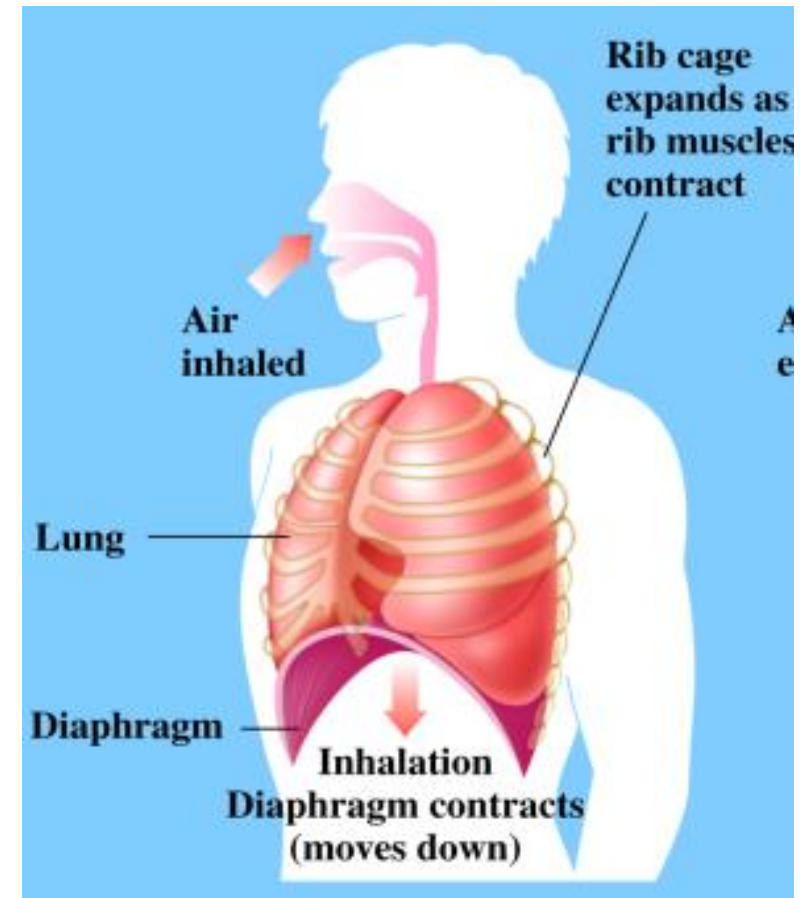
# Boyle's Law



# Boyle's Law

During an inhalation:

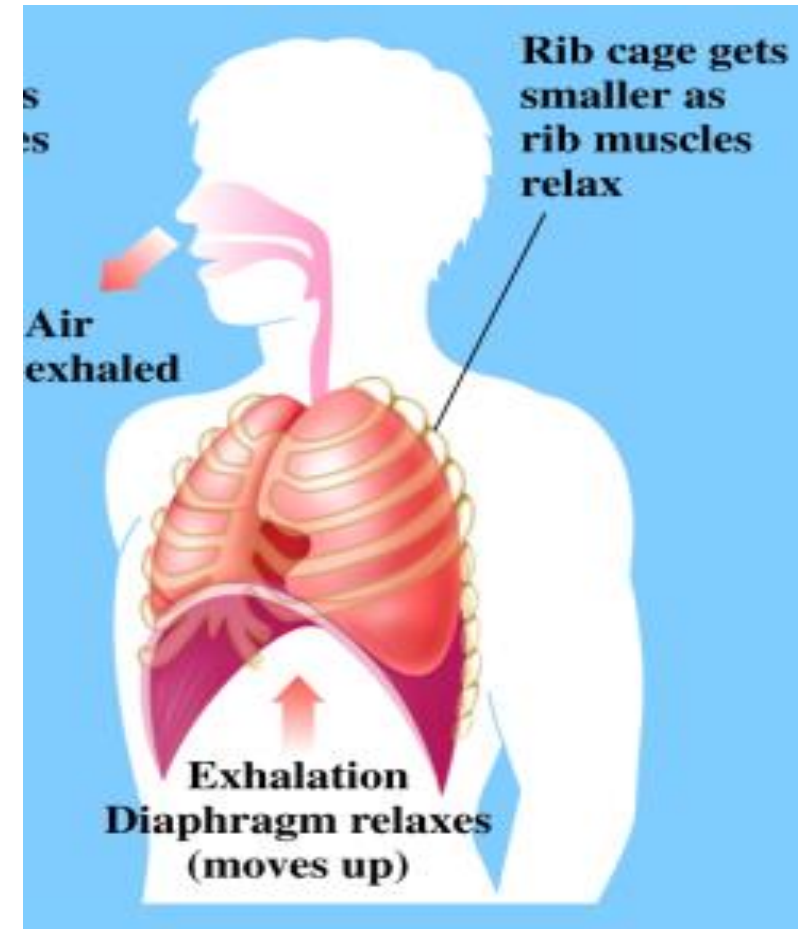
- The lungs expand.
- The pressure in the lungs decreases.
- Air flows towards the lower pressure in the lungs.



# Boyle's Law

During an exhalation:

- The lungs volume decreases.
- The pressure in the lungs increases.
- Air flows from the higher pressure in the lungs to the outside.



# Charles Law

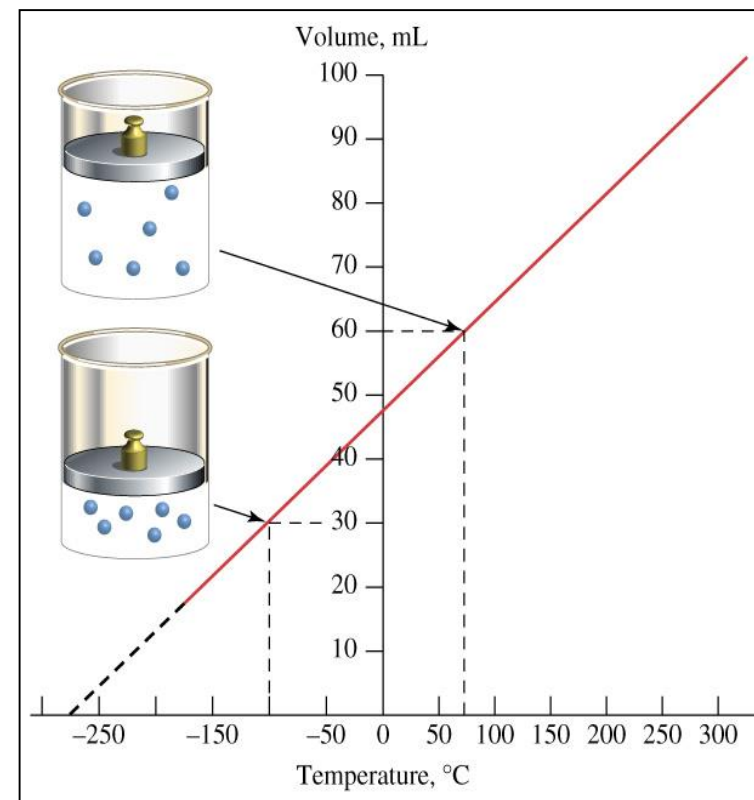
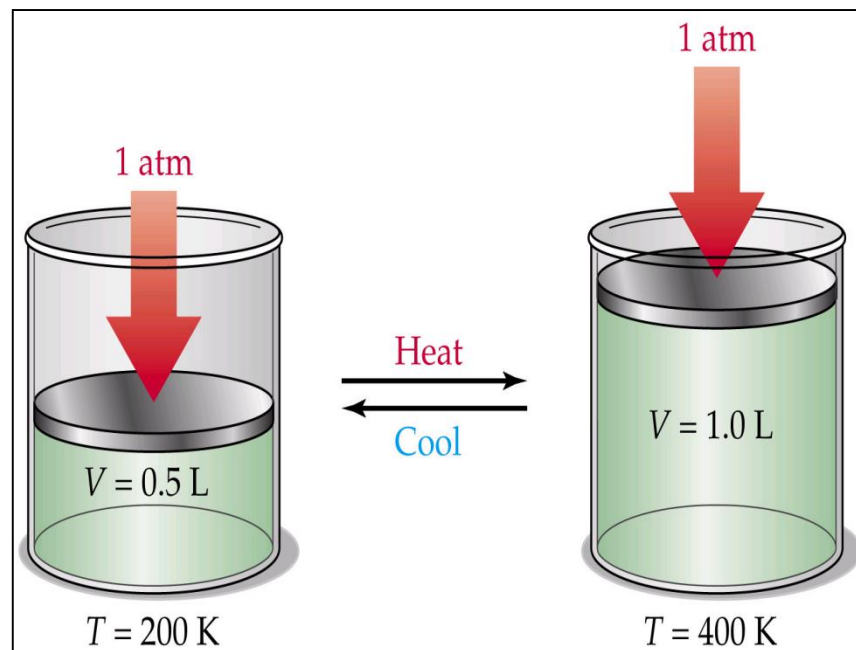
Volume of a gas varies directly with the absolute temperature at constant pressure.

$$V \propto T$$

$$\frac{V}{T} = \text{constant}$$

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

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





# Avogadro's Law

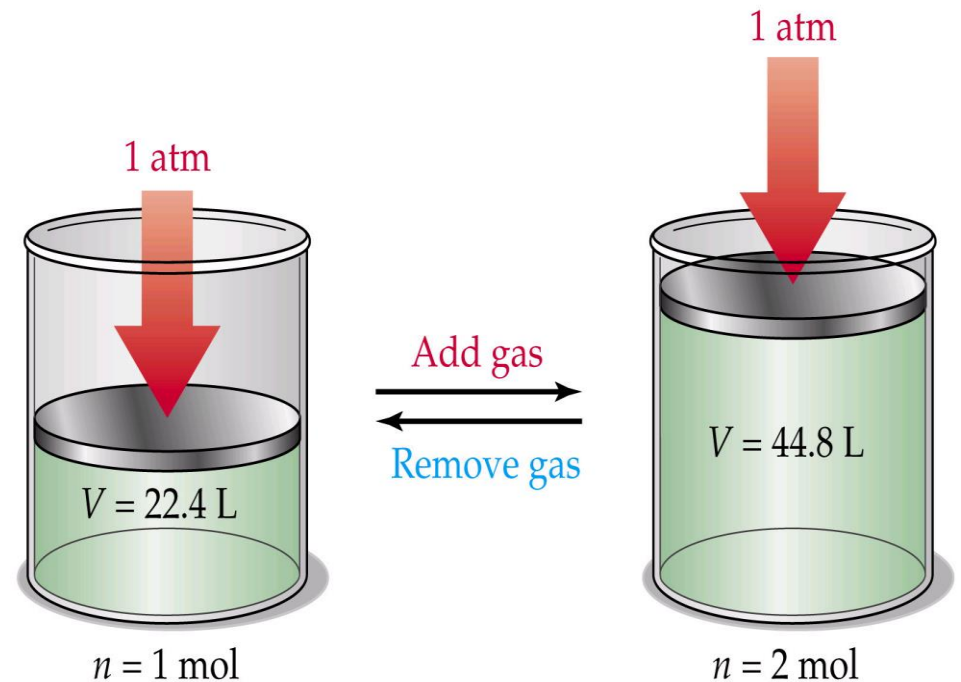
At constant temperature and pressure, the volume of a gas is directly related to the number of moles.

$$V \propto n$$

$$\frac{V}{n} = \text{constant}$$

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

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



# Gay-Lussac's Law

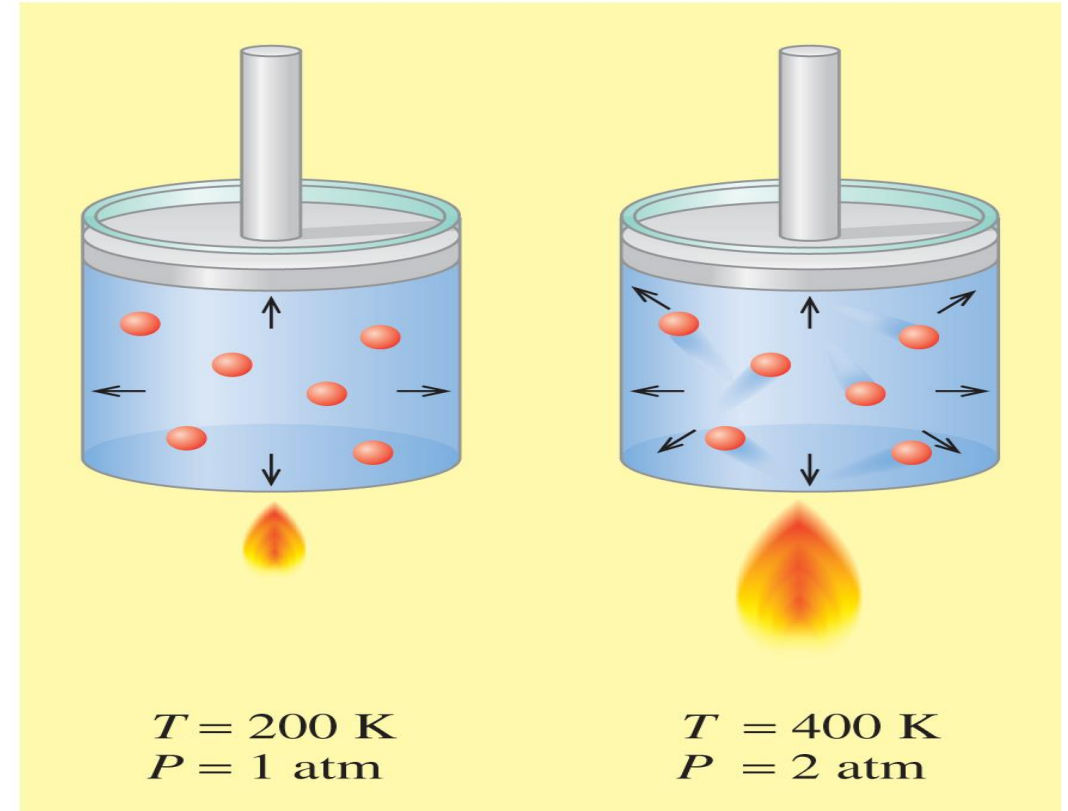
At constant volume, pressure and absolute temperature of a gas are directly related.

$$P \propto T$$

$$\frac{P}{T} = \text{constant}$$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

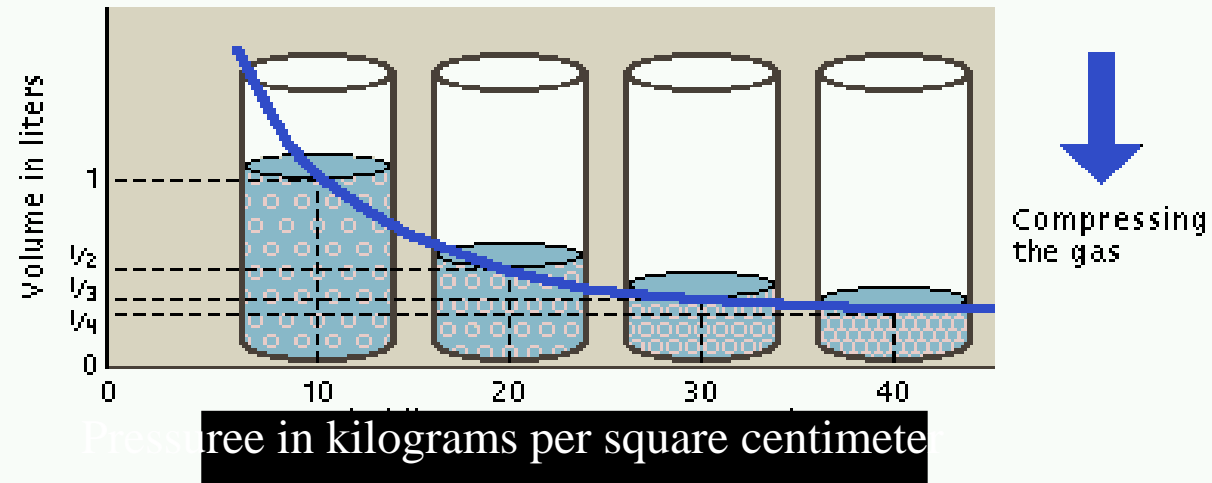
$$\frac{P_1}{P_2} = \frac{T_1}{T_2}$$



# Combined Law

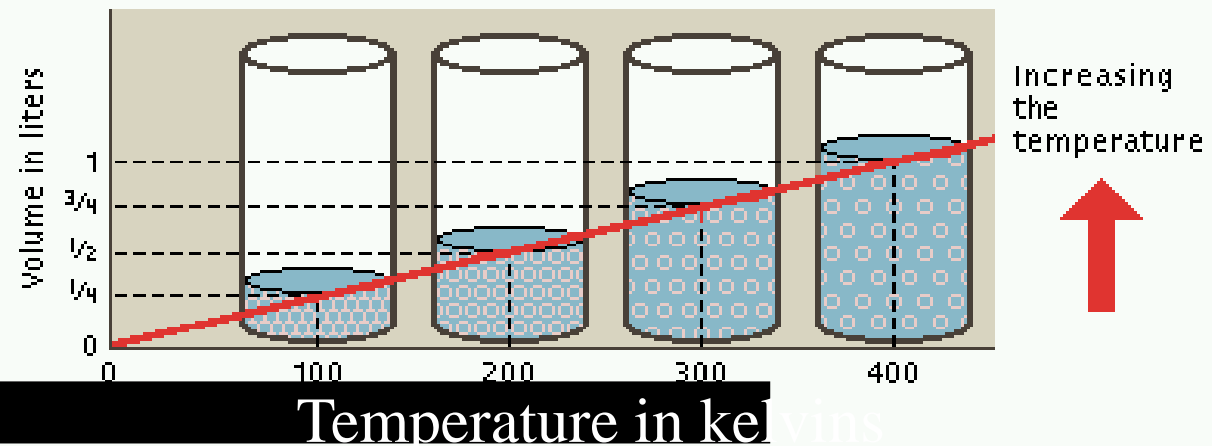
## Boyle's Law

If a gas is held at a **constant temperature**, the volume is inversely proportional to the pressure. Compressing a gas to half of its initial volume doubles its pressure.



## Charles' Law

If a gas is held at a **constant pressure**, the volume is directly proportional to the absolute temperature. Heating a gas to double its original temperature doubles its volume.



# Combined Law

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Boyle's Law     $V = \underline{k}P^{-1}$      $P_1V_1 = P_2V_2$   
(at a fixed temperature)

Charles' Law     $V = bT$      $V_1 / V_2 = T_1 / T_2$   
(at a fixed pressure)

Avogadro     $V = an$     (at a fixed pressure  
and temperature)  
 $n = \text{number of moles}$

$V = nRT P^{-1}$      **$PV = nRT$**     an empirical  
ideal gas law    law

# Ideal vs. Real Gases

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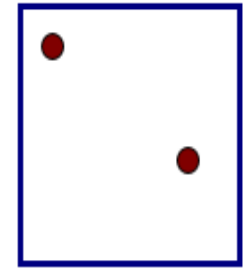
	<b>Ideal Gas</b>	<b>Real Gas</b>
<b>Obey <math>PV=nRT</math></b>	<b>Always</b>	<b>Only at very low P and high T</b>
<b>Molecular volume</b>	<b>Zero</b>	<b>Small but nonzero</b>
<b>Molecular attractions</b>	<b>Zero</b>	<b>Small</b>
<b>Molecular repulsions</b>	<b>Zero</b>	<b>Small</b>

# Real Gases

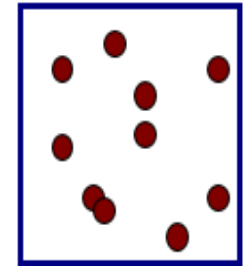
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- Real molecules do take up space and do interact with each other (especially polar molecules).
- Need to add correction factors to the ideal gas law to account for these.
- Ideally, the volume of the molecules was neglected

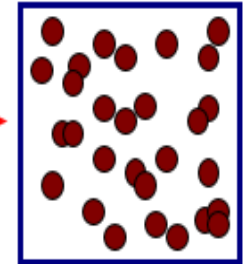
at 1 Atmosphere Pressure



at 10 Atmospheres Pressure



at 30 Atmospheres Pressure



# Real Gases

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But since real gases do have volume, we need:

## Volume Correction

- The actual volume free to move in is less because of particle size.
- More molecules will have more effect.
- Corrected volume  $V' = V - nb$
- “ $b$ ” is a constant that differs for each gas.

# Real Gases

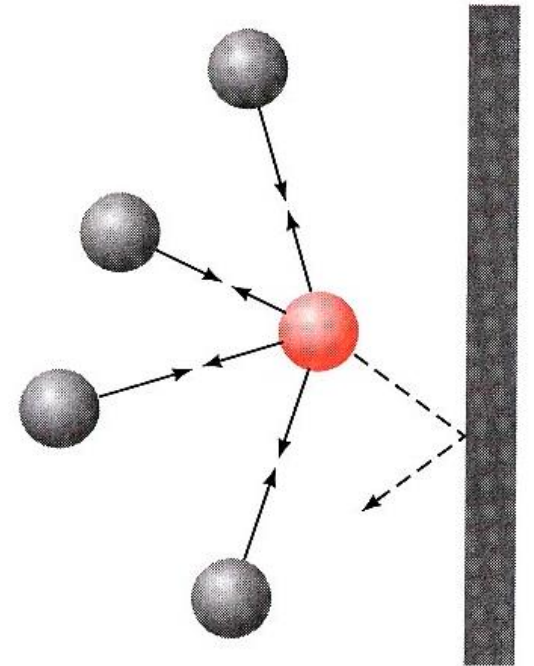
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But since real gases do have volume, we need:

## Pressure Correction

- Because the molecules are attracted to each other, the pressure on the container will be less than ideal.
- Pressure depends on the number of molecules per liter.
- Since two molecules interact, the effect must be squared

$$P_{\text{observed}} = P - a \left(\frac{n}{V}\right)^2$$





# Van Der Waal's Equation

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$$\left[ P_{\text{obs}} + a \left( \frac{n}{V} \right)^2 \right] (V - nb) = nRT$$

Corrected Pressure      Corrected Volume

- “a” and “b” are determined by experiment
- “a” and “b” are different for each gas
- bigger molecules have larger “b”
- “a” depends on both size and polarity

