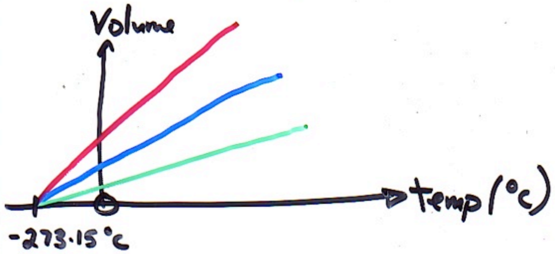
Charles' Law

Volume $\propto$ Temperature (constant pressure + moles of gas)



Absolute temperature scale!

$T(K) = t(°c) + 273.15$

ex: $-273.15°c \longrightarrow 0\ K$ ← Kelvin

ex: $25.00°c \longrightarrow 298.15\ K$

$V \propto T$

$\dfrac{V\uparrow \quad T\uparrow}{V\downarrow \quad T\downarrow}$

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

ex Air at -10°c and a volume of 0.13mL air is heated up to 837°c. What will its volume become?

$V_1 = 0.13\text{mL}$ $\qquad V_2 = ?$

$T_1 = -10 + 273.15 = 263K$ $\qquad T_2 = 837 + 273.15 = 1110K$

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

$\Rightarrow V_2 = \dfrac{V_1 \cdot T_2}{T_1}$

$= \dfrac{0.13\text{mL} \times 1110\cancel{K}}{263\cancel{K}}$

$= 0.55\text{mL}$ ($\approx$ 4 times as large as before)

ex: We have a balloon of helium which has a volume of 2.0L @ 25°c. Q. How cold do we have to make it for the volume to be 0.30L ?

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ (V_1 = 2.0L, V_2 = 0.30L, T_1 = 25 + 273.15 = 298K, T_2 = ?)

$\Rightarrow T_2 = \dfrac{V_2 \times T_1}{V_1}$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow \frac{V_1 T_2^{?}}{V_1} = \frac{T_1 V_2}{V_1} \Rightarrow T_2 = \frac{T_1 V_2}{V_1}$$

$$\Rightarrow T_2 = \frac{0.30\,L \times 298\,K}{2.0\,L} = 45\ K$$

$$T(K) = t(°C) + 273.15$$

$$t(°C) = T(K) - 273.15$$
$$= 45 - 273.15$$
$$= \underline{\underline{-228\,°C}}$$

Avogadro's Law

$V \propto n$ ← # mol gas

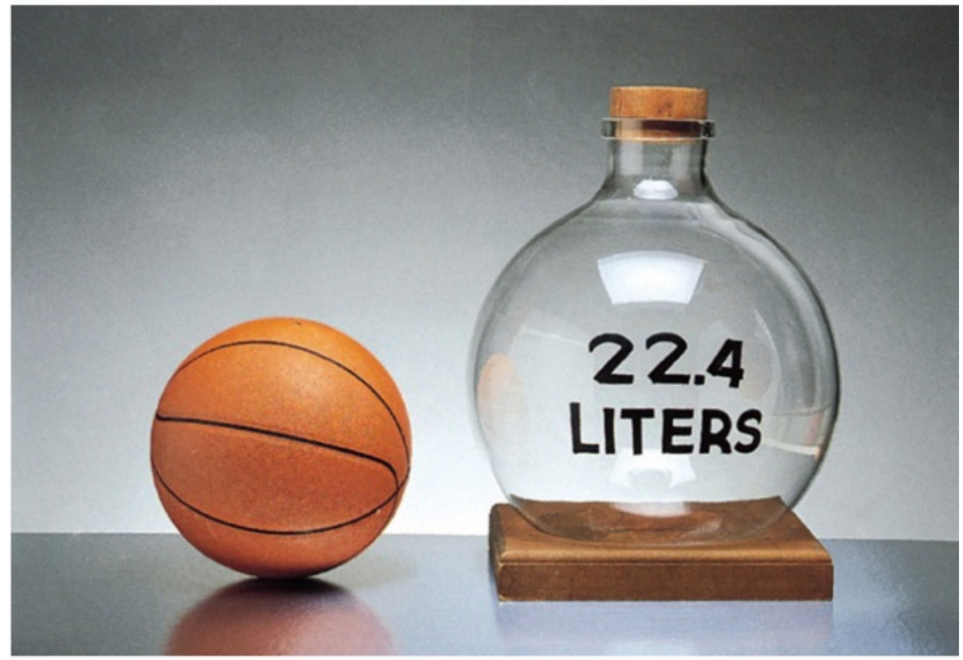
1 mol of gas @ STP occupies 22.4 L

Standard Temperature + Pressure

0 °C and 1 atm

273.15 K   760 mmHg

101,325 Pa



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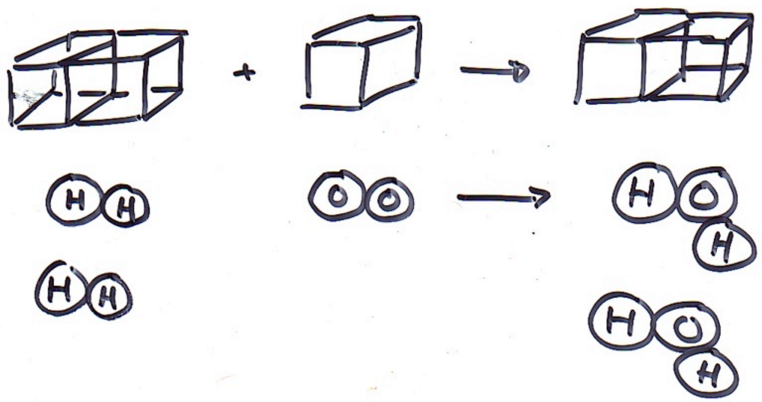
Hydrogen + Oxygen → Water



A B C

Avogadro's $V \propto n$

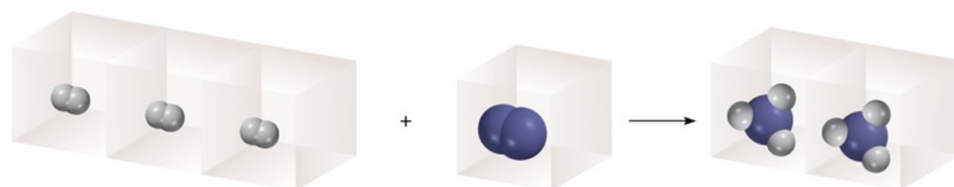
Boyle's $P \propto 1/V$

Charles' $V \propto T$

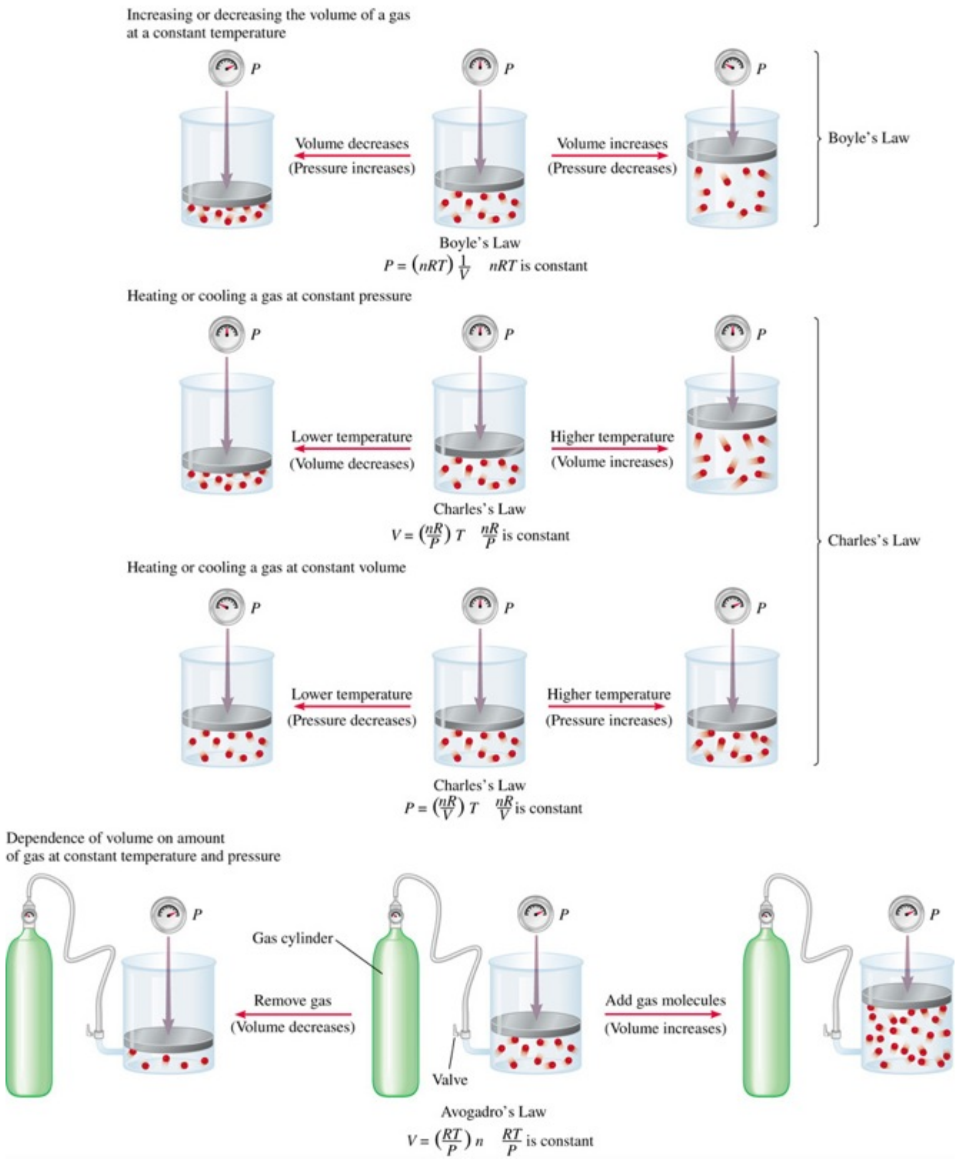
Ideal Gas Equation

$$pV = nRT$$

- p: pressure
- V: volume
- n: #mol gas
- R: Ideal Gas Constant
- T: Temperature (K)



$3H_2(g)$	+	$N_2(g)$	→	$2NH_3(g)$
3 molecules	+	1 molecule	→	2 molecules
3 moles	+	1 mole	→	2 moles
3 volumes	+	1 volume	→	2 volumes



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

ex: What's the pressure of 0.50 mol of $CH_4(g)$ @ −35°C and a volume of 2.10 L?

−35 + 273.15 = 238 K

$pV = nRT \Rightarrow p = \frac{nRT}{V}$

$\Rightarrow p = \frac{0.50\,mol \times 0.08206 \frac{atm \cdot L}{mol \cdot K} \times 238\,K}{2.10\,L}$

= 4.7 atm

Let's send 0.25 mol of $H_2(g)$ to Mars, where t = 5°C and pressure is 0.038 atm. What volume will it occupy?

$pV = nRT \Rightarrow V = \frac{nRT}{P} =$

$$V = \frac{nRT}{P} = \frac{0.25\,\cancel{mol} \times 0.08206\,\frac{\cancel{atm}\cdot L}{\cancel{mol}\cdot\cancel{K}} \times 278\,\cancel{K}}{0.038\,\cancel{atm}}$$

$$= 150\ L$$

$$pV = nRT \longrightarrow R = \frac{pV}{nT}$$

ex: @ STP, V = 22.4 L for 1 mol gas.

$$R = \frac{\overset{\infty}{1\,atm} \times \overset{3sf}{22.4\,L}}{\underset{\infty}{1\,mol} \times \underset{\infty}{273.15\,K}} = 0.0820\,\frac{atm \times L}{mol \times K}$$

$$= \frac{760\,mmHg \times 22.4\,L}{1\,mol \times 273.15\,K} = 62.3\,\frac{mmHg\cdot L}{mol\cdot K}$$

$$= \frac{760\,mmHg \times 22{,}400\,mL}{1\,mol \times 273.15\,K} = 6.23 \times 10^4\,\frac{mmHg\cdot mL}{mol\cdot K}$$

$$\frac{pV}{nT} = R$$

$$\Rightarrow \boxed{\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}}$$

normally, # mol gas doesn't change (unless we have a leak!)

$$\boxed{\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}}$$

Combined Gas Law.