W-GC General Chemistry 3/23

Ideal Gas Law = $PV = nRT$

Boyle's Law

$P_1V_1 = P_2V_2$

$$\frac{P_1V_1}{V_2} = \frac{P_2V_2}{V_2}$$

1.) Gas 3.5 atm and 15 L →
transfer 20L container.

Pressure?

$$P_2 = \frac{P_1V_1}{V_2}$$

$$\frac{(3.5\,atm)(15\,\cancel{L})}{20\,\cancel{L}}$$

2.6 atm

$$PV = nRT$$

$$\frac{PV}{nRT} = \frac{nRT}{nRT}$$

$$\frac{PV}{nRT} = 1$$

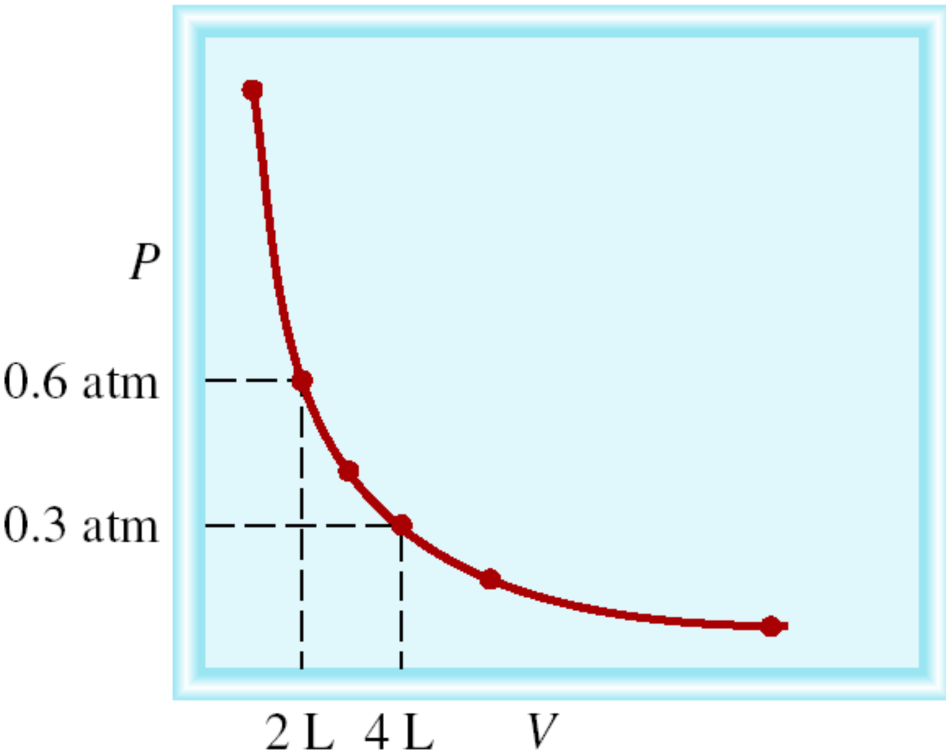
$$\frac{P_1V_1}{n_1RT_1} = \frac{P_2V_2}{n_2RT_2}$$

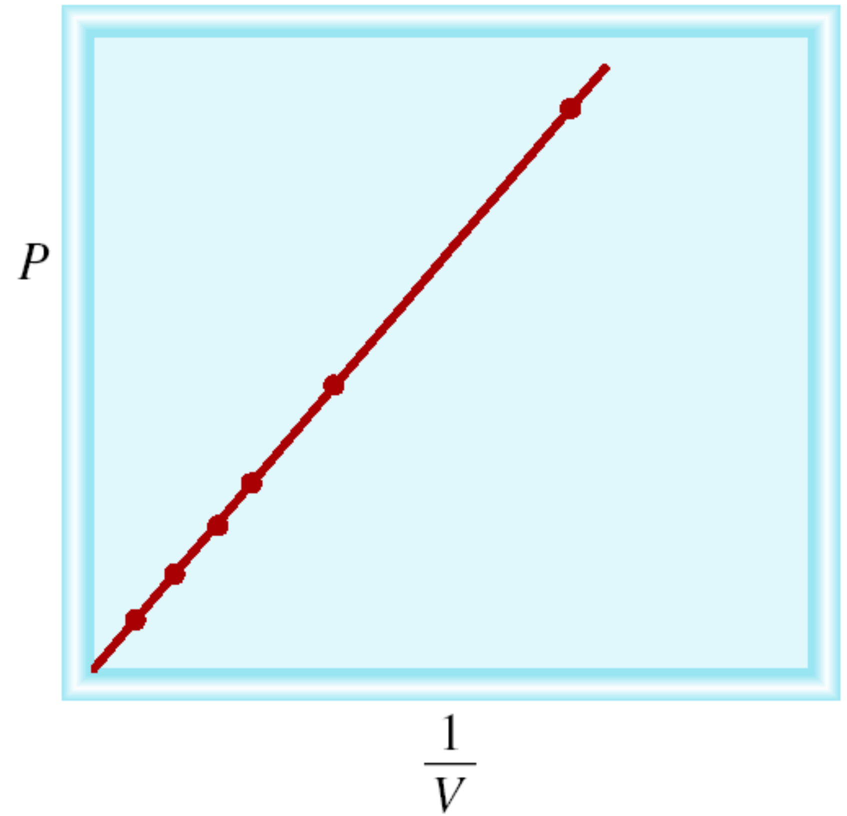
Boyle's

$$\frac{P_1V_1}{\cancel{n_1RT_1}} = \frac{P_2V_2}{\cancel{n_2RT_2}}$$

n, T = constant
hold temp and amount constant

Boyle's Law





$P \alpha 1/V$

$P \times V = \text{constant}$

$P_1 \times V_1 = P_2 \times V_2$

Constant temperature
Constant amount of gas

A sample of chlorine gas occupies a volume of 946 mL at a pressure of 726 mmHg. What is the pressure of the gas (in mmHg) if the volume is reduced at constant temperature to 154 mL?

[Ideal Gas Law $\boxed{PV} = \boxed{nRT}$]

$$P \times V = constant$$

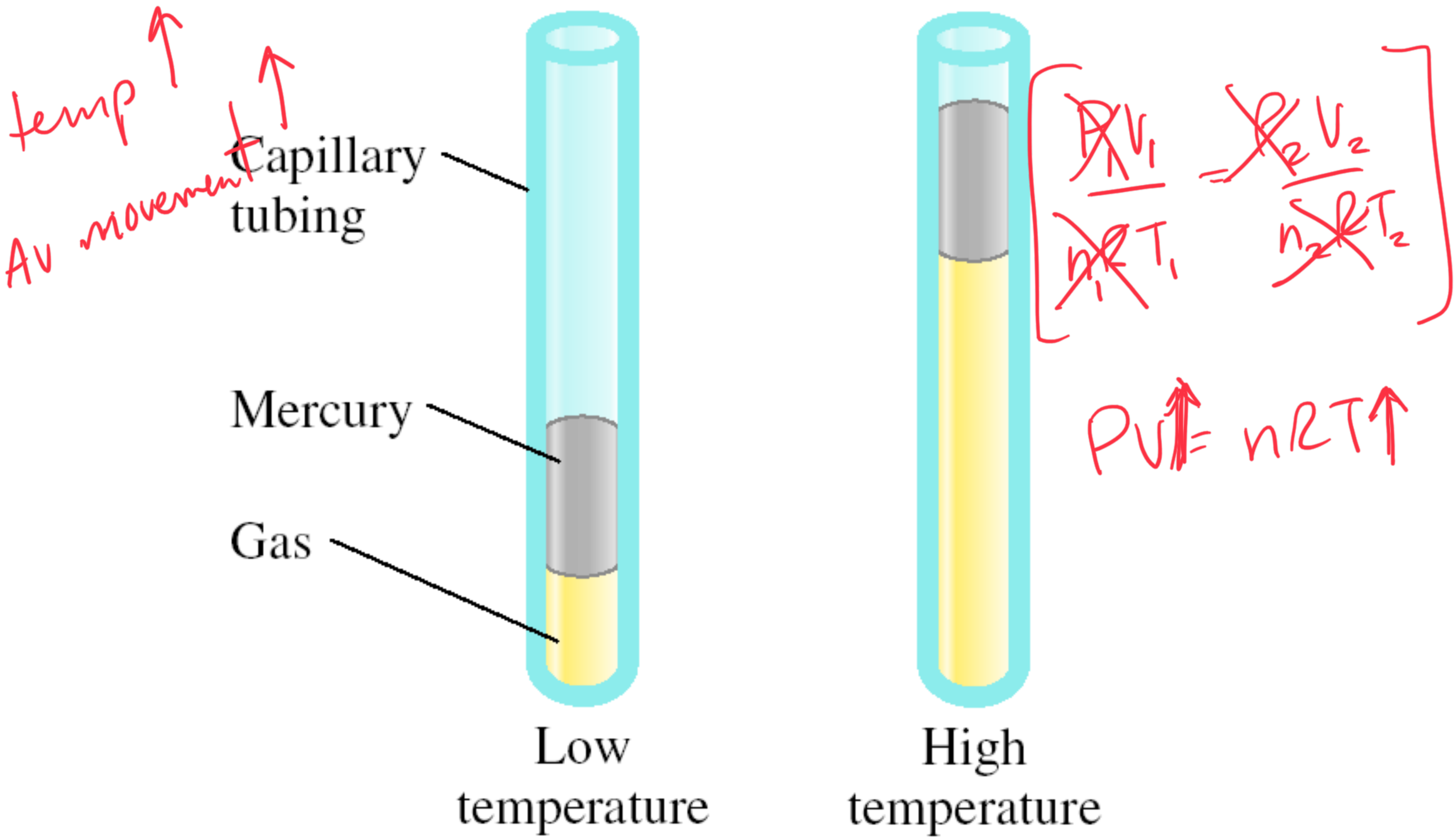
$$P_1 \times V_1 = P_2 \times V_2$$

$P_1 = 726$ mmHg $\quad P_2 = ?$

$V_1 = 946$ mL $\quad V_2 = 154$ mL

$$P_2 = \frac{P_1 \times V_1}{V_2} = \frac{726 \text{ mmHg} \times 946 \text{ mL}}{154 \text{ mL}} = 4460 \text{ mmHg}$$

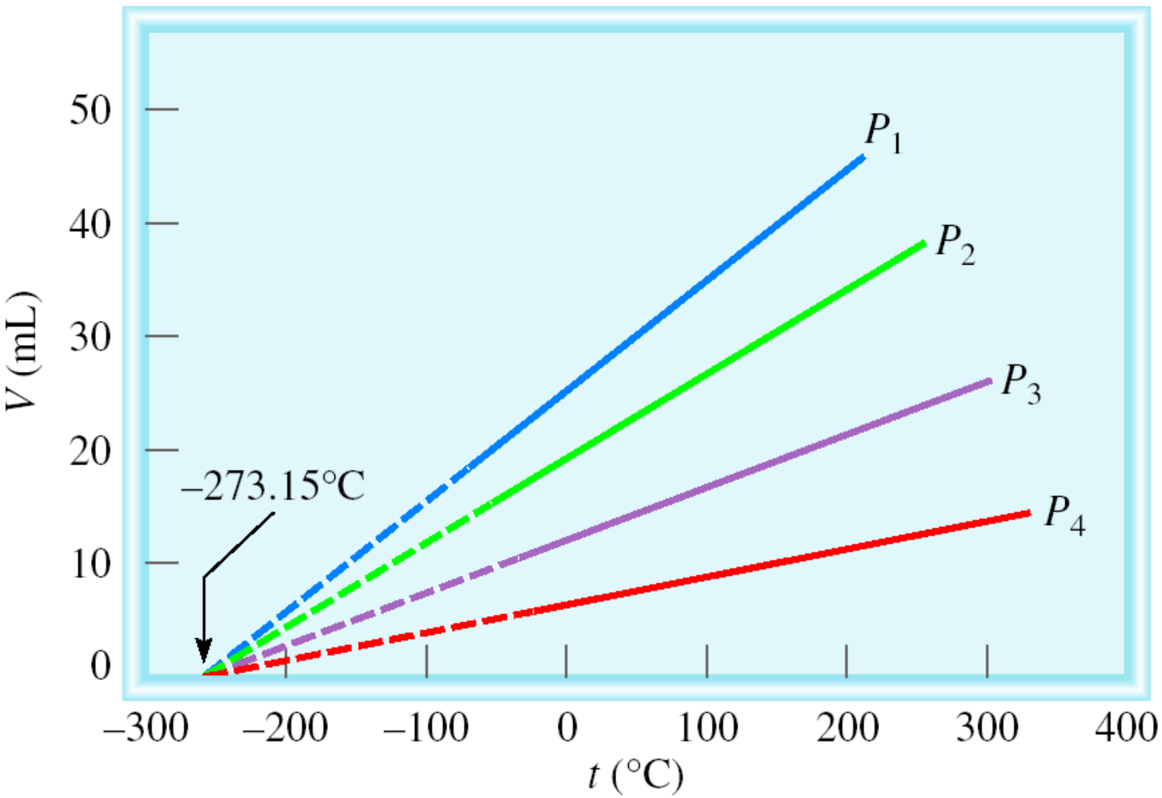
Variation in Gas Volume with Temperature at Constant Pressure



As T increases V increases

Variation of Gas Volume with Temperature at Constant Pressure

As temp ↑
volume ↑



Charles' & Gay-Lussac's Law

$V \propto T$

V = constant x T

$V_1/T_1 = V_2/T_2$

Temperature **must** be in Kelvin

T (K) = t (^{0}C) + 273.15

A sample of carbon monoxide gas occupies 3.20 L at 125 ^{0}C. At what temperature will the gas occupy a volume of 1.54 L if the pressure remains constant?

$$V_1/T_1 = V_2/T_2$$

$V_1 = 3.20$ L $\qquad V_2 = 1.54$ L

$T_1 = 398.15$ K $\qquad T_2 = ?$

$$T_1 = 125\ (^{0}\text{C}) + 273.15\ (\text{K}) = 398.15\ \text{K}$$

$$T_2 = \frac{V_2 \times T_1}{V_1} = \frac{1.54\ \text{L} \times 398.15\ \text{K}}{3.20\ \text{L}} = 192\ \text{K}$$

Charle's Law / Gay-Lussac's

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

$V_1 = 700\,mL$ $T_1 = 20°C$

$T_2 = 100°C$ $V_2 = ?$

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

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

$$V_2 = \frac{(700\,mL)(373\,K)}{293\,K} = 891$$

890 mL

0.89 L

$T_1 = 20 + 273 = 293\,K$

$T_2 = 100 + 273 = 373\,K$

Avogadro's Law

$V \propto$ number of moles (n)

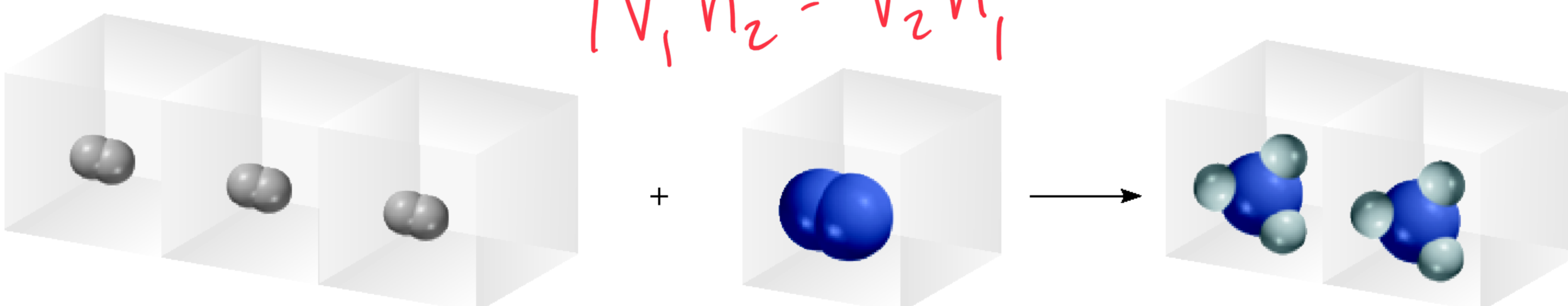
Constant temperature
Constant pressure

V = constant x n

$V_1 / n_1 = V_2 / n_2$

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

$V_1 n_2 = V_2 n_1$



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

Ammonia burns in oxygen to form nitric oxide (NO) and water vapor. How many volumes of NO are obtained from one volume of ammonia at the same temperature and pressure?

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

$2H_2O \rightarrow 2H_2 + O_2$

$$4NH_3 + 5O_2 \longrightarrow 4NO + 6H_2O$$

$$1 \text{ mole } NH_3 \longrightarrow 1 \text{ mole } NO$$

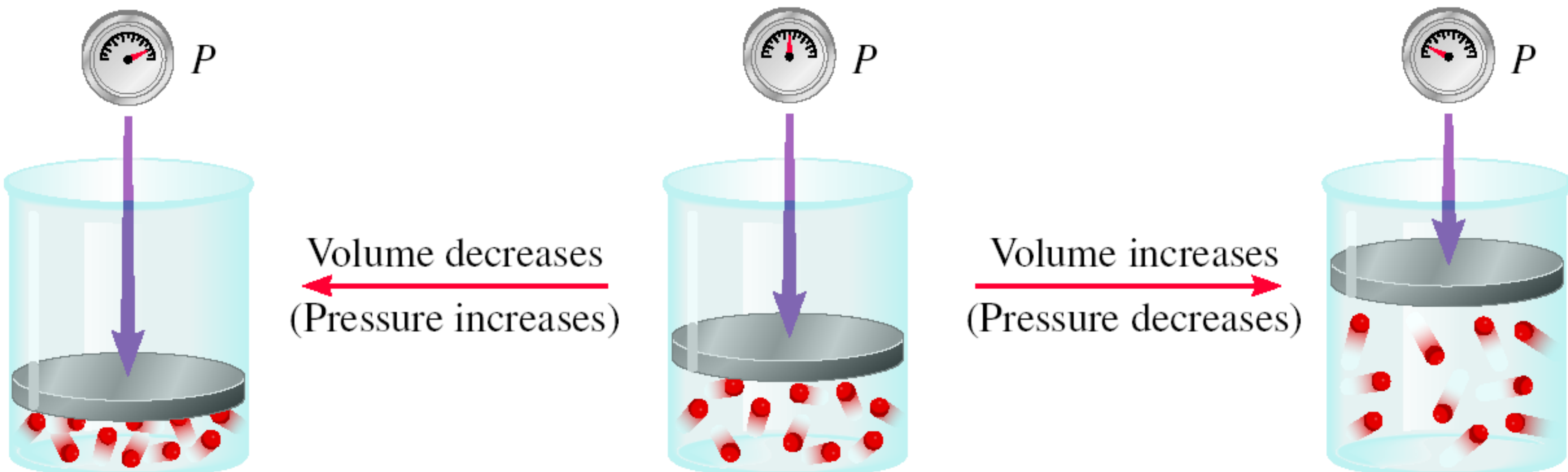
At constant *T* and *P*

$$1 \text{ volume } NH_3 \longrightarrow 1 \text{ volume } NO$$

Summary of Gas Laws

Boyle's Law

Increasing or decreasing the volume of a gas at a constant temperature

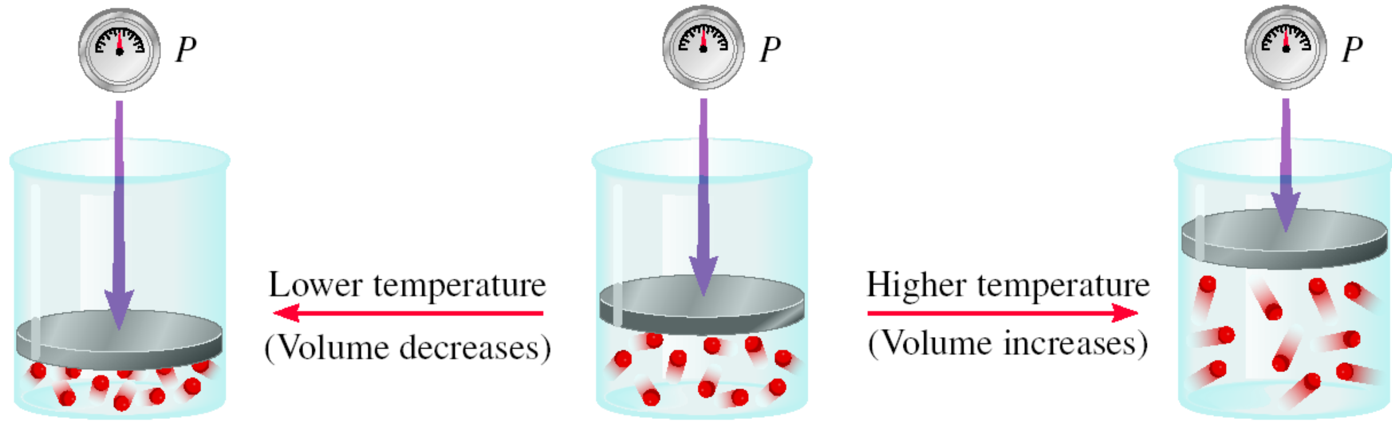


Boyle's Law

$$P = (nRT)\frac{1}{V} \quad nRT \text{ is constant}$$

Charles Law

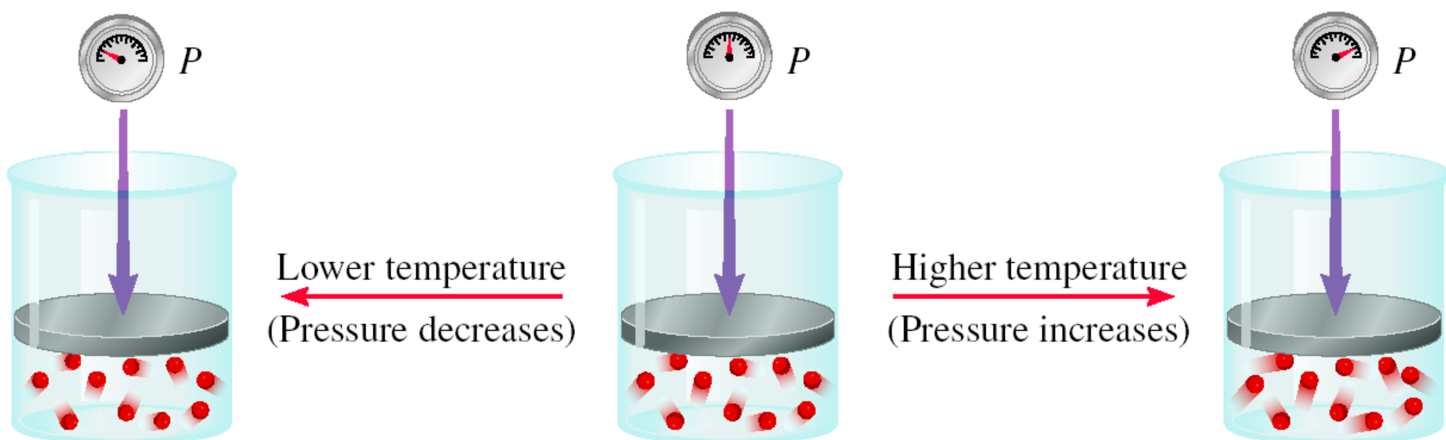
Heating or cooling a gas at constant pressure



Charles's Law

$$V = \left(\frac{nR}{P}\right) T \quad \frac{nR}{P} \text{ is constant}$$

Heating or cooling a gas at constant volume

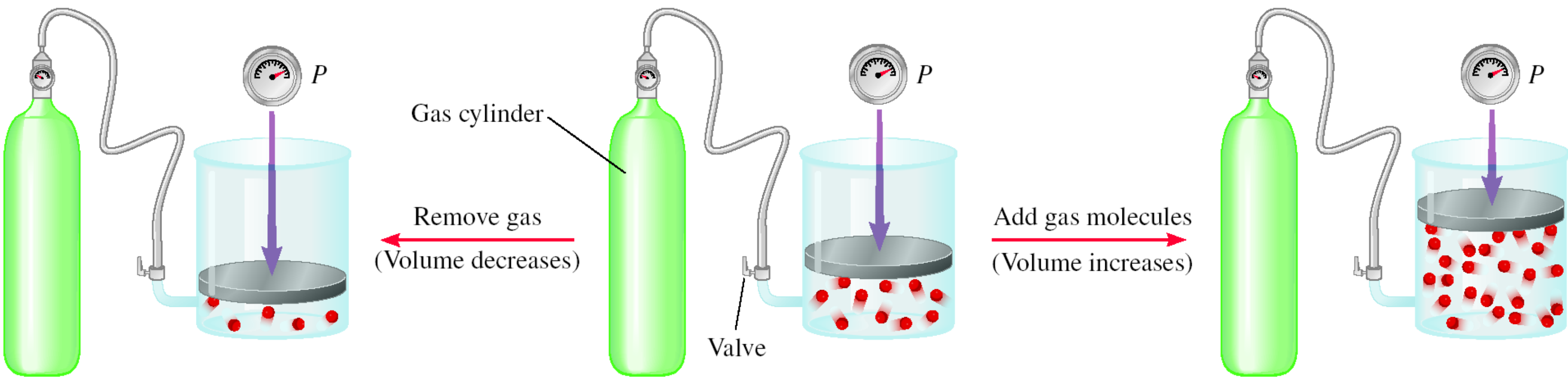


Charles's Law

$$P = \left(\frac{nR}{V}\right) T \quad \frac{nR}{V} \text{ is constant}$$

Avogadro's Law

Dependence of volume on amount of gas at constant temperature and pressure



Avogadro's Law

$$V = \left(\frac{RT}{P}\right) n \qquad \frac{RT}{P} \text{ is constant}$$

Ideal Gas Equation

Boyle's law: $P \propto \frac{1}{V}$ (at constant n and T)

Charles' law: $V \propto T$ (at constant n and P)

Avogadro's law: $V \propto n$ (at constant P and T)

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

$$V = \text{constant} \times \frac{nT}{P} = R\frac{nT}{P}$$

R is the **gas constant**

$$PV = nRT$$

The conditions 0 ^{0}C and 1 atm are called **standard temperature and pressure (STP).**

Experiments show that at STP, 1 mole of an ideal gas occupies 22.414 L.

$$PV = nRT$$

$$R = \frac{PV}{nT} = \frac{(1\ \text{atm})(22.414\text{L})}{(1\ \text{mol})(273.15\ \text{K})}$$

$$R = 0.082057\ \text{L} \bullet \text{atm} / (\text{mol} \bullet \text{K})$$

What is the volume (in liters) occupied by 49.8 g of HCl at STP?

$T = 0\ ^0C = 273.15\ K$

$P = 1\ atm$

$$PV = nRT$$

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

$$n = 49.8\ g \times \frac{1\ mol\ HCl}{36.45\ g\ HCl} = 1.37\ mol$$

$$V = \frac{1.37\ mol \times 0.0821\ \frac{L \bullet atm}{mol \bullet K} \times 273.15\ K}{1\ atm}$$

$$V = 30.7\ L$$

Argon is an inert gas used in lightbulbs to retard the vaporization of the filament. A certain lightbulb containing argon at 1.20 atm and 18 ^{0}C is heated to 85 ^{0}C at constant volume. What is the final pressure of argon in the lightbulb (in atm)?

$$PV = nRT \qquad n,\ V \text{ and } R \textbf{ are constant}$$

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

$P_1 = 1.20$ atm $\qquad P_2 = ?$

$T_1 = 291$ K $\qquad T_2 = 358$ K

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

$$P_2 = P_1 \, x \, \frac{T_2}{T_1} = 1.20 \text{ atm x } \frac{358 \text{ K}}{291 \text{ K}} = 1.48 \text{ atm}$$