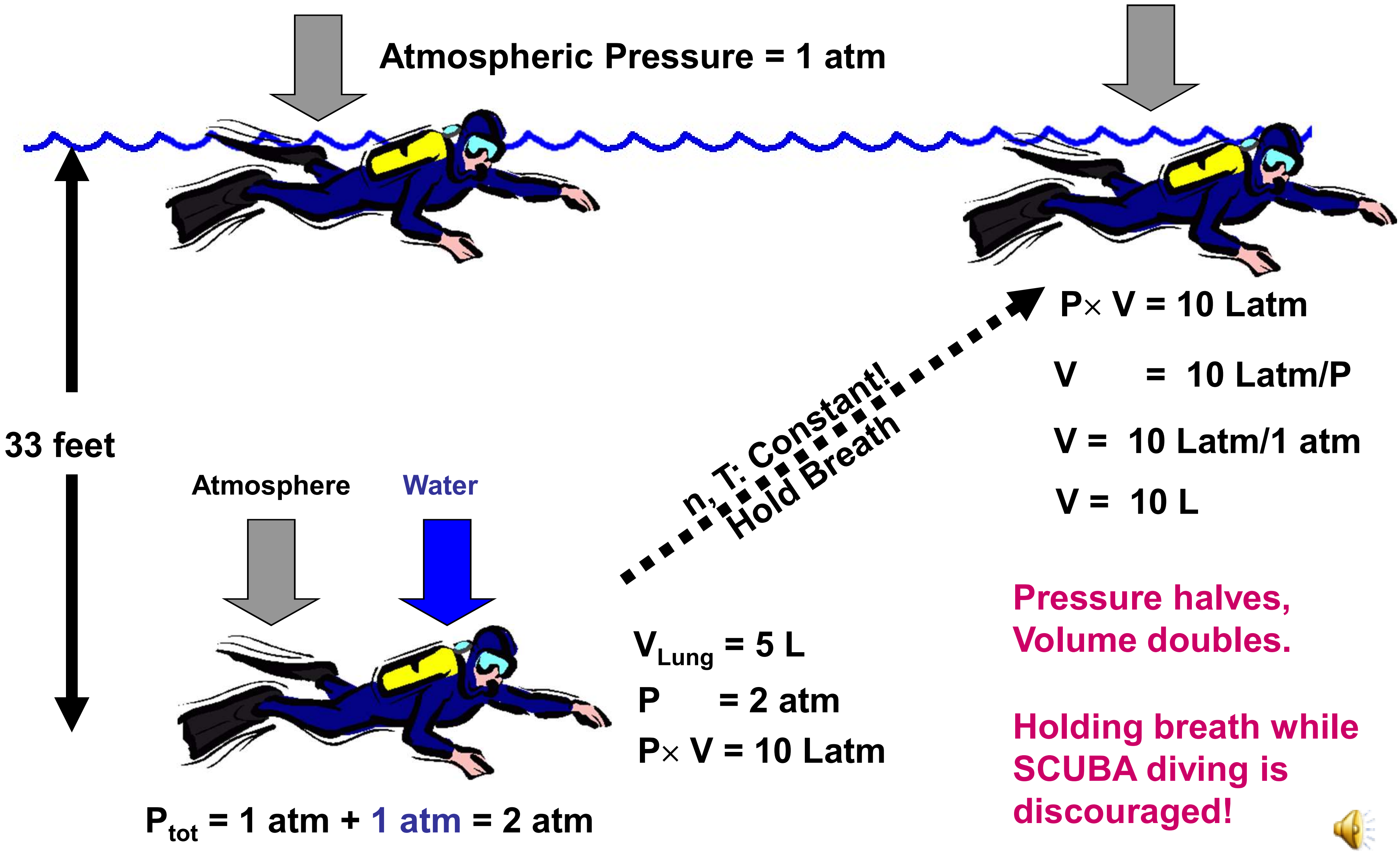
Gas Law Relationships (P, V, n, T)

Name	Constants	Relationship	Equation
Charle's Law	(P, n)	$V \propto T$ Volume is directly proportional to temperature	$\frac{V_i}{T_i} = k = \frac{V_f}{T_f}$
Boyle's Law	(T, n)	$P \propto \frac{1}{V}$ Pressure is inversely proportional to temperature	$P_i \times V_i = k = P_f \times V_f$
Avagadro's Law	(T, P)	$V \propto n$ Volume is directly proportional to the number of moles of gas.	$\frac{V_i}{n_i} = k = \frac{V_f}{n_f}$
Combined Gas Law	(n)	$P \times V \propto T$ The product of pressure and volume is proportional to temperature.	$\frac{P_i V_i}{T_i} = k = \frac{P_f V_f}{T_f}$

Boyles law (PV = k)



Avagadro's Law

(Molar Volume)

$$\frac{V}{n} = k$$

The ratio of volume to moles of particles is constant assuming pressure and temperature don't change

Identity of the gas doesn't matter!

N_2, Ar, O_2, Ne, He, CO_2 etc…

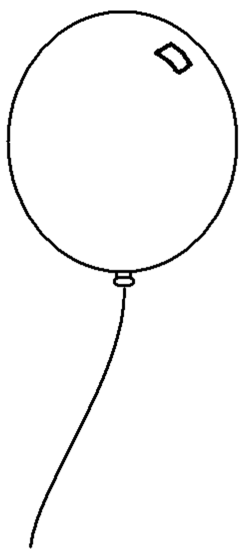
Standard Temperature & Pressure

S.T.P.

Temperature = 0°C (273.15 K)

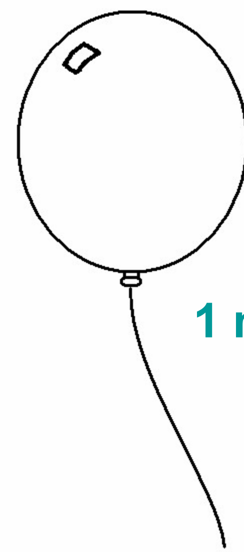
Pressure = 1.00 atm

1 mole $He_{(g)}$



Volume = 22.414 Liters

1 mole $N_{2(g)}$



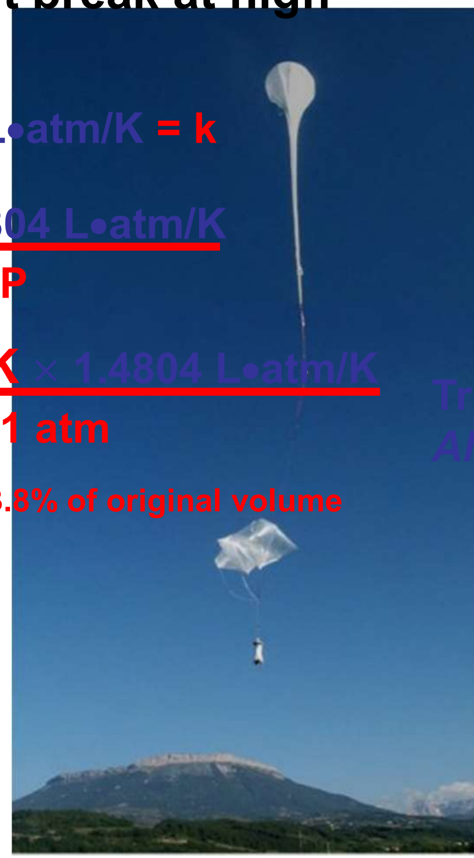
Volume = 22.414 Liters

Molar volume of an ideal gas @ STP: 1 mole = 22.414 Liters

Combined Gas Law (PV/T = k)

What is the max. volume of gas that can be used so that the balloon won't break at high altitude?

Decrease T: balloon contracts
Decrease P: balloon expands



PV/T = 1.4804 L•atm/K = k

$$V = \frac{T \times 1.4804\ L\bullet atm/K}{P}$$

Launch:

P = 1.00 atm
T = 25.0 °C (298.15 K)
V = ? L

$$V = \frac{298.15K \times 1.4804\ L\bullet atm/K}{1\ atm}$$

V = 441 L 13.8% of original volume

Troposphere:
Altitude = 10 miles

P = 0.10 atm
T = -57.0 °C 9 (216.15K)
$V_{balloon}$ = 3200 L (max)

PV/T = k

PV/T = (0.10atm)(3200L)/216.15K)

PV/T = (0.10atm)(3200L)/216.15K)

PV/T = 1.4804 L•atm/K = k

Using Molar Volume

Problem 5.54

When 35.6 L of ammonia and 40.5 L of oxygen gas at **STP** burn, nitrogen monoxide and water are produced. After products return to STP, how many grams of nitrogen monoxide are present?

Excess		*Limiting*		30.01	
$4\ NH_{3(g)}$	+	$5\ O_{2(g)}$	$\rightarrow$	$4\ NO_{(g)}$	$6\ H_2O_{(g)}$
35.6 L		**40.5 L**			

1.59 **moles** (from NH_3) → **1.59** **moles** (NO)

1.81 **moles** (from O_2) → **1.44** **moles** (NO)

STP

Pressure = 1 atm.

Temp = 0°C

$$\frac{\mathbf{35.6\ L}}{\mathbf{1}} \times \frac{\mathbf{1\ mole_{NH3}}}{\mathbf{22.414\ L}} \times \frac{4\ \text{moles}_{\ NO}}{4\ \text{moles}_{\ NH3}} = \mathbf{1.59\ mole_{NO}}$$

$$\frac{\mathbf{40.5\ L}}{\mathbf{1}} \times \frac{\mathbf{1\ mole_{O2}}}{\mathbf{22.414\ L}} \times \frac{4\ \text{moles}_{\ NO}}{5\ \text{moles}_{\ O2}} = \mathbf{1.44\ mole_{NO}} \times \frac{30.01\ g_{\ NO}}{1\ \text{moles}_{\ NO}} = 43.4\ g_{\ NO}$$

Ideal Gas Law

$$\frac{P_i V_i}{n_i T_i} = \textit{constant} = \frac{P_f V_f}{n_f T_f}$$

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

R: Universal Gas Law Constant

R = 8.31447 J/(mol · K)

R = 0.0820578 (L ·atm)/(mol · K)

Let units determine the correct gas law constant to use.

Ideal Gas Law (cont.)

$$\mathbf{P\ V = n\ \mathit{R}\ T}$$

- P: Pressure *(atm)*
- V: Volume (L)
- n: moles
- *R*: Gas Law Constant $\left(\frac{L\ atm}{mole\ K}\right)$
- T: Temperature (K)

$$n = \frac{PV}{RT} \qquad V = \frac{nRT}{P} \qquad P = \frac{nRT}{V}$$