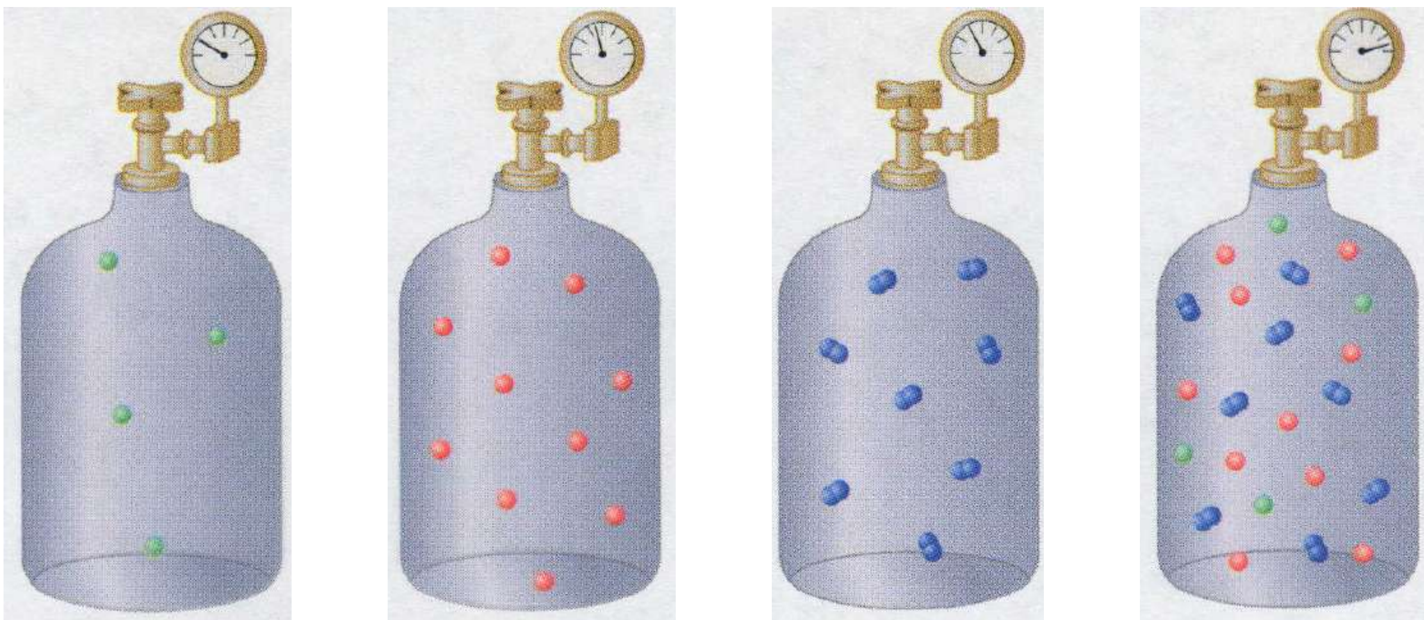
Partial Pressures

200 kPa + 500 kPa + 400 kPa = 1100 kPa ? kPa

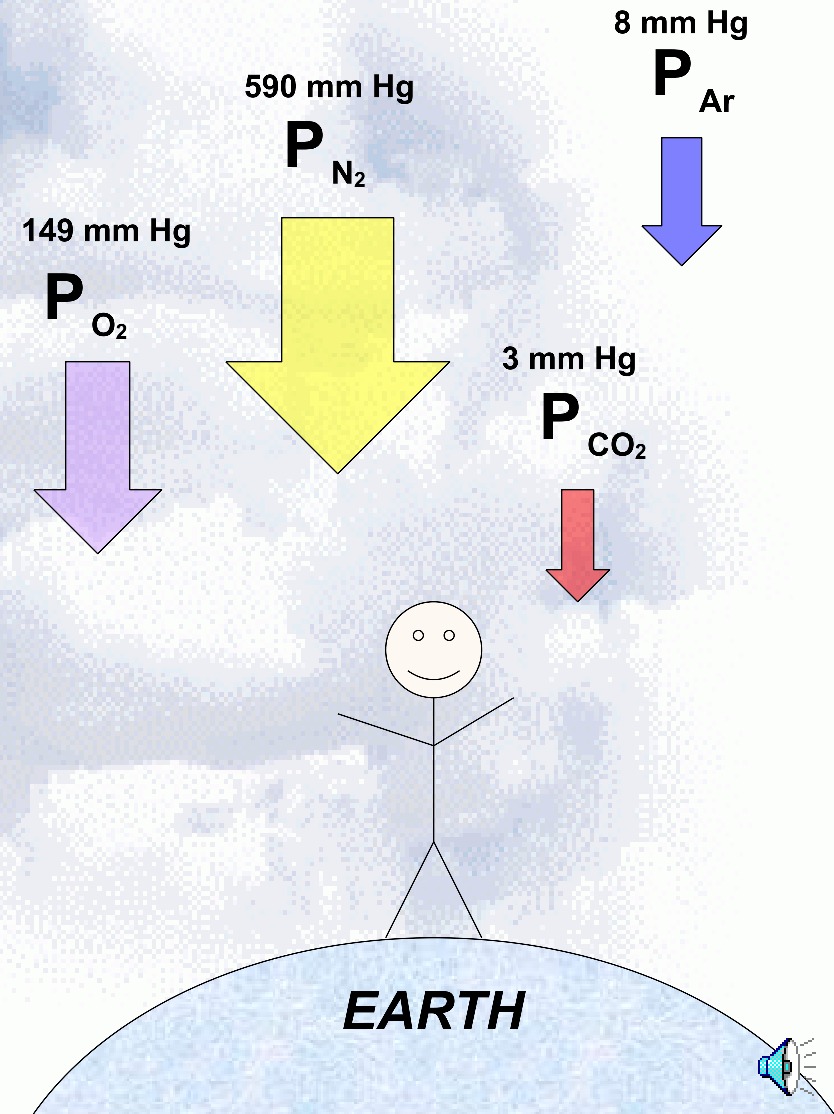


Dalton's Law of Partial Pressures & Air Pressure

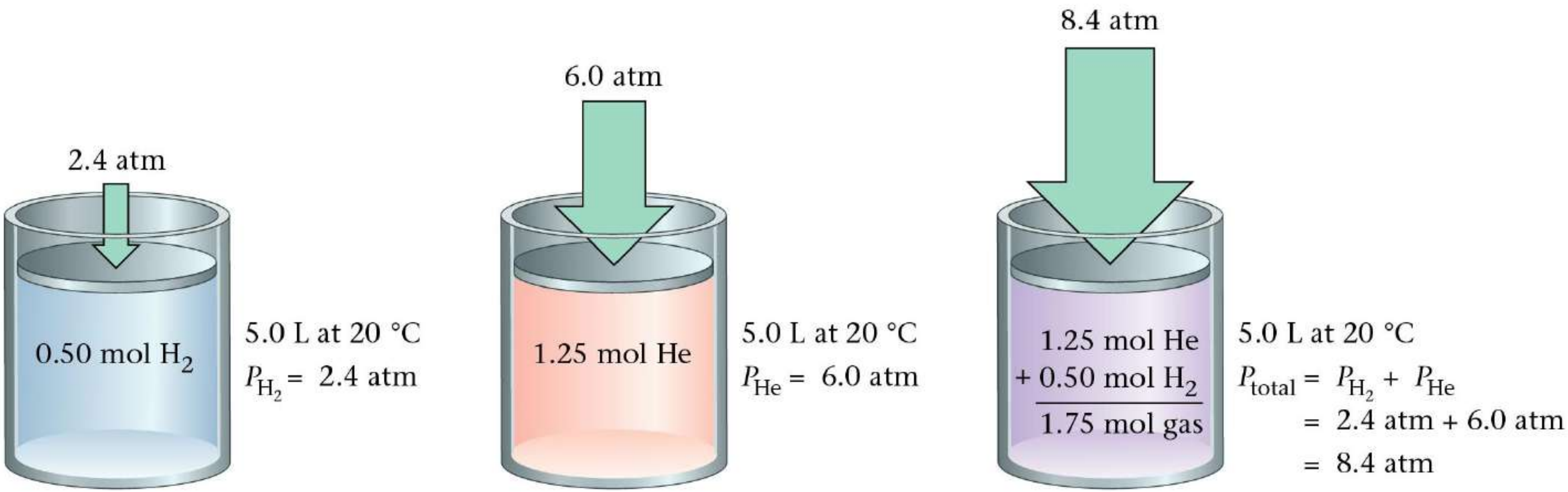
$$P_{Total} = P_{O_2} + P_{N_2} + P_{CO_2} + P_{Ar}$$

$$P_{Total} = 149 + 590 + 3 + 8 \text{ mm Hg}$$

$$P_{Total} = 750 \text{ mm Hg}$$

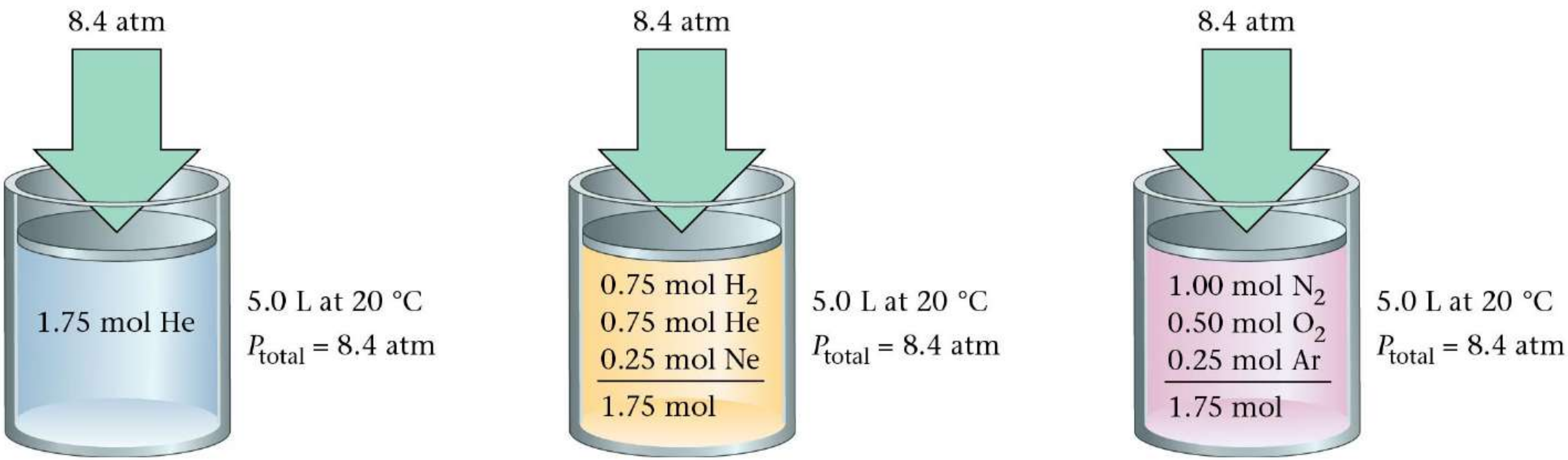


Dalton's Partial Pressures



Zumdahl, Zumdahl, DeCoste, World of Chemistry 2002, page 421

Dalton's Law

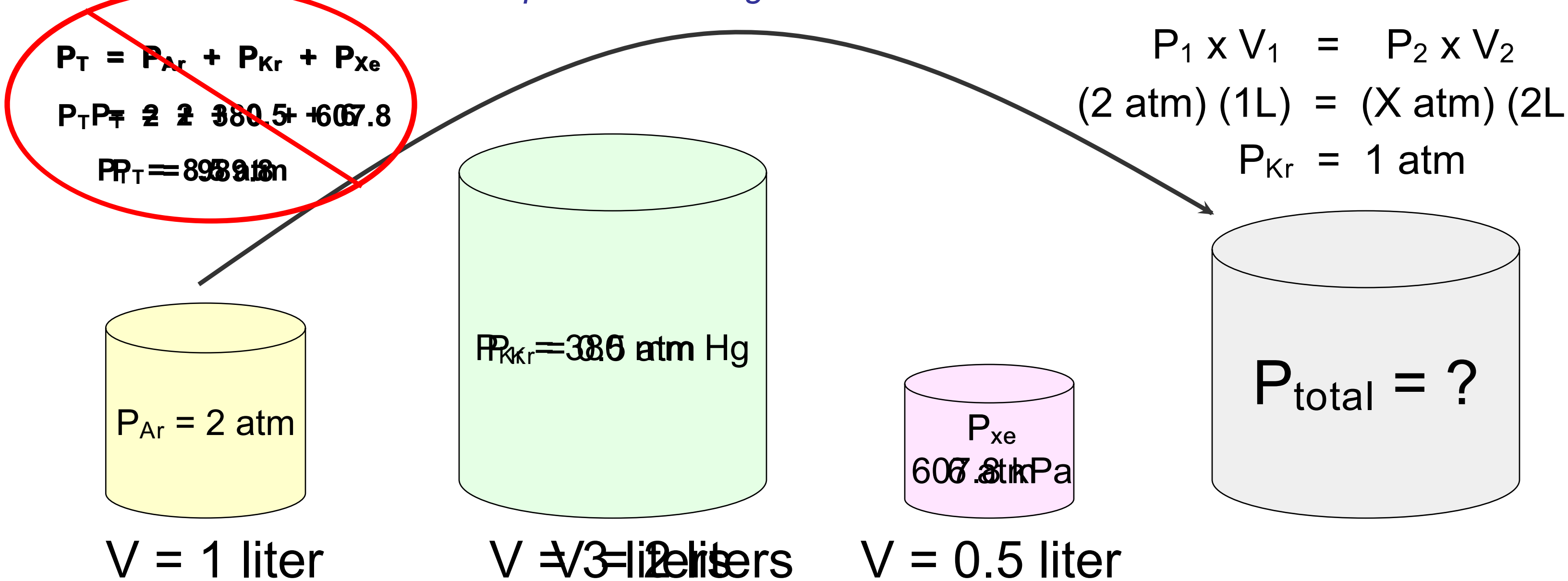


Zumdahl, Zumdahl, DeCoste, World of Chemistry 2002, page 422

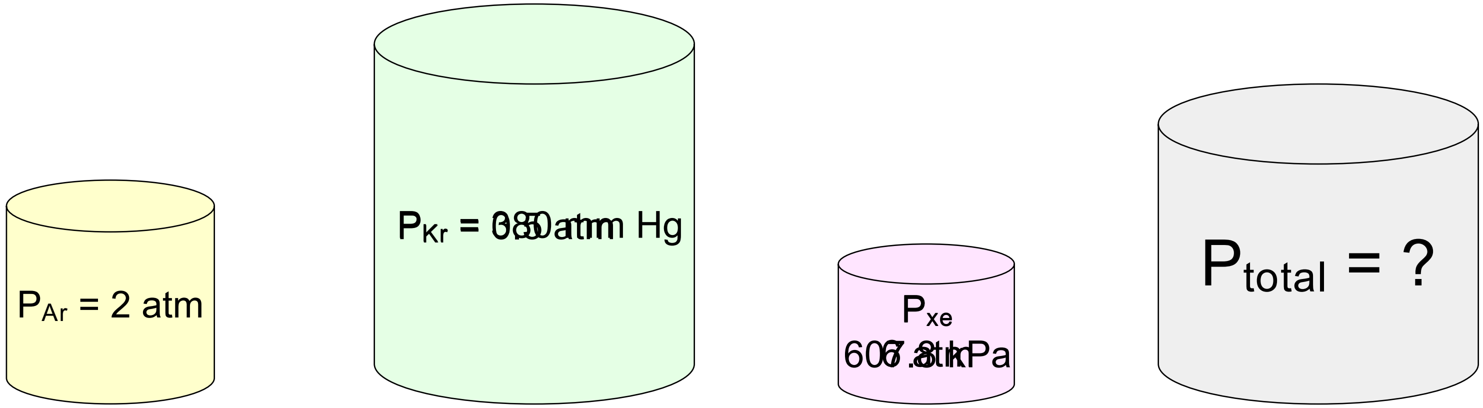
Dalton's Law Applied

Suppose you are given four containers – three filled with noble gases. The first 1 L container is filled with argon and exerts a pressure of 2 atm. The second 3 liter container is filled with krypton and has a pressure of 380 mm Hg. The third 0.5 L container is filled with xenon and has a pressure of 607.8 kPa. If all these gases were transferred into an empty 2 L container…what would be the pressure in the "new" container?

What would the pressure of argon be if transferred to 2 L container?



…just add them up



V = 1 liter V = 3 liters V = 0.5 liter V = 2 liters

Dalton's Law of Partial Pressures

"**Total Pressure = Sum of the Partial Pressures**"

$$P_T = P_{Ar} + P_{Kr} + P_{Xe} + \ldots$$

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

$$(0.5\text{ atm})\,(3\text{L}) = (X\text{ atm})\,(2\text{L})$$

$$P_{Kr} = 0.75\text{ atm}$$

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

$$(6\text{ atm})\,(0.5\text{ L}) = (X\text{ atm})\,(2\text{L})$$

$$P_{xe} = 1.5\text{ atm}$$

$$P_T = 1\text{ atm} + 0.75\text{ atm} + 1.5\text{ atm}$$

$$P_T = 3.25\text{ atm}$$

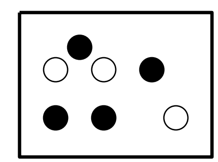
Dalton's Law of Partial Pressures

In a gaseous mixture, a gas's **partial pressure** is the one the gas would exert if it were by itself in the container.

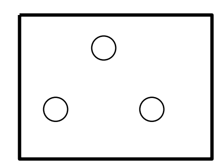
The mole ratio in a mixture of gases determines each gas's partial pressure.

Total pressure of mixture (3.0 mol He and 4.0 mol Ne) is 97.4 kPa.

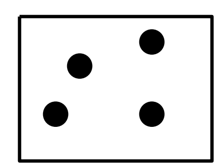
Find partial pressure of each gas



$$P_{He} = \frac{3 \text{ mol He}}{7 \text{ mol gas}} (97.4 \text{ kPa}) = \boxed{41.7 \text{ kPa}}$$

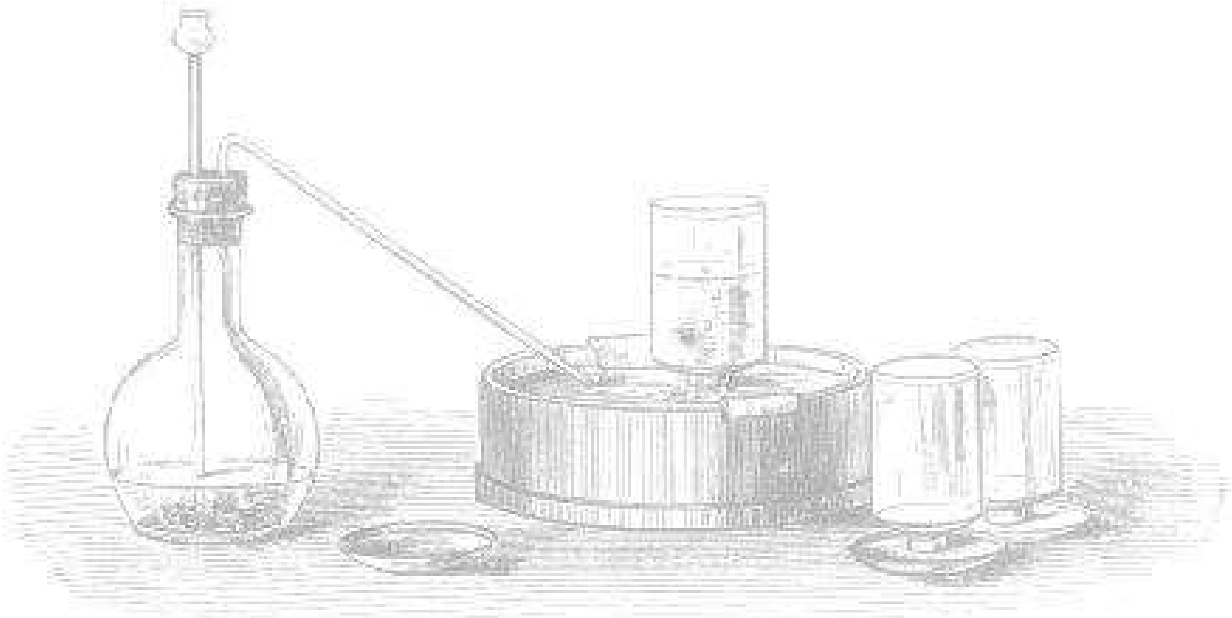


$$P_{Ne} = \frac{4 \text{ mol Ne}}{7 \text{ mol gas}} (97.4 \text{ kPa}) = \boxed{55.7 \text{ kPa}}$$



Dalton's Law:
the total pressure exerted by a mixture of gases is the sum of all the partial pressures

$$P_Z = P_{A,Z} + P_{B,Z} + \ldots$$



80.0 g each of He, Ne, and Ar are in a container.
The total pressure is 780 mm Hg.
Find each gas's partial pressure.

$$80\text{ g He}\left(\frac{1\text{ mol}}{4\text{ g}}\right) = 20\text{ mol He}$$

$$80\text{ g Ne}\left(\frac{1\text{ mol}}{20\text{ g}}\right) = 4\text{ mol Ne}$$

$$80\text{ g Ar}\left(\frac{1\text{ mol}}{40\text{ g}}\right) = 2\text{ mol Ar}$$

Total: 26 mol gas

P_{He} = $^{20}/_{26}$ of total

P_{Ne} = $^{4}/_{26}$ of total

P_{Ar} = $^{2}/_{26}$ of total

$$P_{He} = 600\text{ mm Hg},\ P_{Ne} = 120\text{ mm Hg},\ P_{Ar} = 60\text{ mm Hg}$$

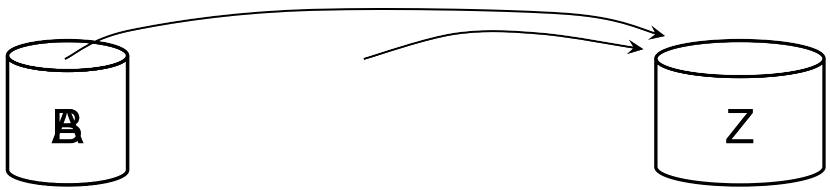
Dalton's Law: $P_Z = P_{A,Z} + P_{B,Z} + \ldots$

Two 1.0 L containers, A and B, contain gases under 2.0 and 4.0 atm, respectively. Both gases are forced into Container B. Find total pres. of mixture in B.

	P_X	V_X	V_Z	$P_{X,Z}$
A	2.0 atm	1.0 L	1.0 L	2.0 atm
B	4.0 atm	1.0 L		4.0 atm

Total = 6.0 atm

Two 1.0 L containers, A and B, contain gases under 2.0 and 4.0 atm, respectively. Both gases are forced into Container Z (^{w}/vol. 2.0 L). Find total pres. of mixture in Z.



	P_X	V_X	V_Z	$P_{X,Z}$
A	2.0 atm	1.0 L	2.0 L	1.0 atm
B	4.0 atm	1.0 L		2.0 atm

Total = 3.0 atm

$$P_A V_A = P_Z V_Z$$

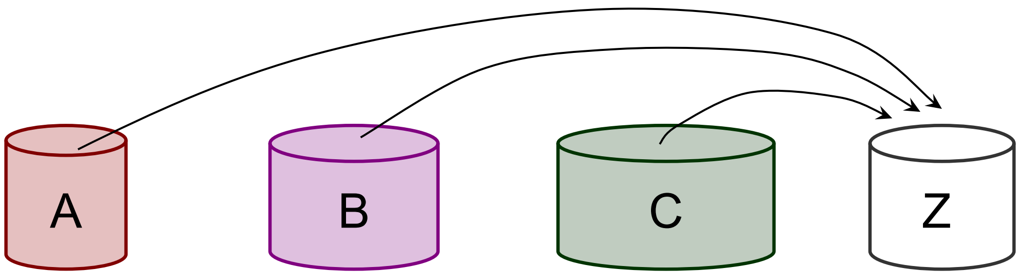
2.0 atm (1.0 L) = X atm (2.0 L)

X = 1.0 atm

$$P_B V_B = P_Z V_Z$$

4.0 atm (1.0 L) = X atm (2.0 L)

Find total pressure of mixture in Container Z.



1.3 L 2.6 L 3.8 L 2.3 L

3.2 atm 1.4 atm 2.7 atm X atm

	P_X	V_X	V_Z	$P_{X,Z}$
A	3.2 atm	1.3 L		1.8 atm
B	1.4 atm	2.6 L	2.3 L	1.6 atm
C	2.7 atm	3.8 L		4.5 atm

Total = 7.9 atm

$$P_A V_A = P_Z V_Z$$

3.2 atm (1.3 L) = X atm (2.3 L)

X = 1.8 atm

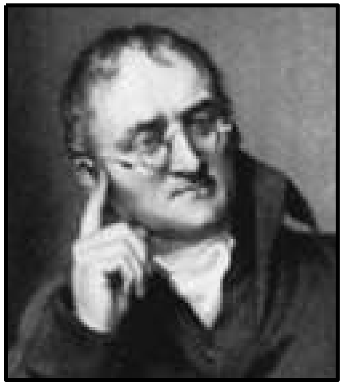
$$P_B V_B = P_Z V_Z$$

1.4 atm (2.6 L) = X atm (2.3 L)

$$P_C V_C = P_Z V_Z$$

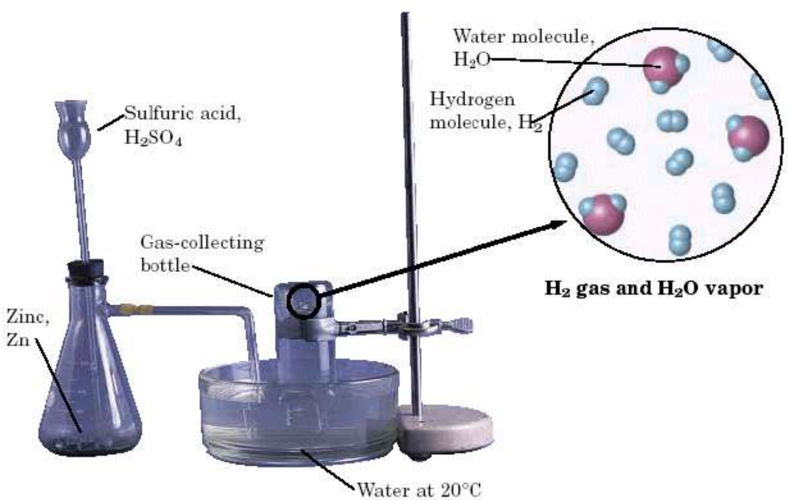
2.7 atm (3.8 L) = X atm (2.3 L)

Dalton's Law



The total pressure of a mixture of gases equals the sum of the partial pressures of the individual gases.

$$P_{total} = P_1 + P_2 + \ldots$$



When a H_2 gas is collected by water displacement, the gas in the collection bottle is actually a mixture of H_2 and water vapor.

Courtesy Christy Johannesson www.nisd.net/communicationsarts/pages/chem

Dalton's Law

Hydrogen gas is collected over water at 22.5°C. Find the pressure of the dry gas if the atmospheric pressure is 94.4 kPa.

The total pressure in the collection bottle is equal to atmospheric pressure and is a mixture of H_2 and water vapor.

GIVEN:	WORK:
P_{H_2} = ?	$P_{total} = P_{H_2} + P_{H_2O}$
P_{total} = 94.4 kPa	$94.4\ kPa = P_{H_2} + 2.72\ kPa$
P_{H_2O} = 2.72 kPa	$P_{H_2} = 91.7\ kPa$

Look up water-vapor pressure on p.899 for 22.5°C.

Sig Figs: Round to least number of decimal places.

Dalton's Law

A gas is collected over water at a temp of 35.0°C when the barometric pressure is 742.0 torr. What is the partial pressure of the dry gas?

The total pressure in the collection bottle is equal to barometric pressure and is a mixture of the "gas" and water vapor.

GIVEN:	WORK:
$P_{gas} = ?$	$P_{total} = P_{gas} + P_{H_2O}$
$P_{total} = 742.0\text{ torr}$	$742.0\text{ torr} = P_{H_2} + 42.2\text{ torr}$
$P_{H_2O} = 42.2\text{ torr}$	$P_{gas} = 699.8\text{ torr}$

Look up water-vapor pressure on p.899 for 35.0°C.

Sig Figs: Round to least number of decimal places.

Dalton's Law of Partial Pressures

1. Container A (with volume 1.23 dm^3) contains a gas under 3.24 atm of pressure. Container B (with volume 0.93 dm^3) contains a gas under 2.82 atm of pressure. Container C (with volume 1.42 dm^3) contains a gas under 1.21 atm of pressure. If all of these gases are put into Container D (with volume 1.51 dm^3), what is the pressure in Container D?

	P_x	V_x	P_D	V_D
A	3.24 atm	1.23 dm^3	2.64 atm	1.51 dm^3
B	2.82 atm	0.93 dm^3	1.74 atm	
C	1.21 atm	1.42 dm^3	1.14 atm	
TOTAL			$P_T = P_A + P_B + P_C$ **5.52 atm**	

$$(P_A)(V_A) = (P_D)(V_D)$$
$$(3.24\text{ atm})(1.23\text{ dm}^3) = (x\text{ atm})(1.51\text{ dm}^3)$$
$$(P_A) = 2.64\text{ atm}$$

$$(P_B)(V_B) = (P_D)(V_D)$$
$$(2.82\text{ atm})(0.93\text{ dm}^3) = (x\text{ atm})(1.51\text{ dm}^3)$$
$$(P_B) = 1.74\text{ atm}$$

$$(P_C)(V_A) = (P_D)(V_D)$$
$$(1.21\text{ atm})(1.42\text{ dm}^3) = (x\text{ atm})(1.51\text{ dm}^3)$$
$$(P_C) = 1.14\text{ atm}$$

Dalton's Law of Partial Pressures

3. Container A (with volume 150 mL) contains a gas under an unknown pressure. Container B (with volume 250 mL) contains a gas under 628 mm Hg of pressure. Container C (with volume 350 mL) contains a gas under 437 mm Hg of pressure. If all of these gases are put into Container D (with volume 300 mL), giving it 1439 mm Hg of pressure, find the original pressure of the gas in Container A.

	P_x	V_x	P_D	V_D
A	STEP 4) **P_A**	150 mL	STEP 3) **406 mm Hg**	
B	628 mm Hg	250 mL	STEP 2) **523 mm Hg**	300 mL
C	437 mm Hg	350 mL	STEP 1) **510 mm Hg**	
TOTAL			$P_T = P_A + P_B + P_C$ 1439 mm Hg	

STEP 1)

$(P_C)(V_C) = (P_D)(V_D)$

$(437)(350) = (x)(300)$

$(P_C) = 510$ mm Hg

STEP 2)

$(P_B)(V_B) = (P_D)(V_D)$

$(328)(250) = (x)(300)$

$(P_B) = 523$ mm Hg

STEP 3)

1439
-510
-523
406 mm Hg

STEP 4)

$(P_A)(V_A) = (P_D)(V_D)$

$(P_A)(150 \text{ mL}) = (406 \text{ mm Hg})(300 \text{ mL})$

(P_A) 812 mm Hg

Table of Partial Pressures of Water

Vapor Pressure of Water					
Temperature (°C)	Pressure (kPa)	Temperature (°C)	Pressure (kPa)	Temperature (°C)	Pressure (kPa)
0	0.6	21	2.5	30	4.2
5	0.9	22	2.6	35	5.6
8	1.1	23	2.8	40	7.4
10	1.2	24	3.0	50	12.3
12	1.4	25	3.2	60	19.9
14	1.6	26	3.4	70	31.2
16	1.8	27	3.6	80	47.3
18	2.1	28	3.8	90	70.1
20	2.3	29	4.0	100	101.3

χ

Mole Fraction

The ratio of the number of moles of a given component in a mixture to the total number of moles in the mixture.

$$\chi_2 = \frac{n_2}{n_{total}} = \frac{P_2}{P_{total}}$$

The partial pressure of oxygen was observed to be 156 torr in air with total atmospheric pressure of 743 torr. Calculate the mole fraction of O_2 present.

$$\chi_{O_2} = \frac{P_{O_2}}{P_{total}} = \frac{156\,torr}{743\,torr} = 0.210$$

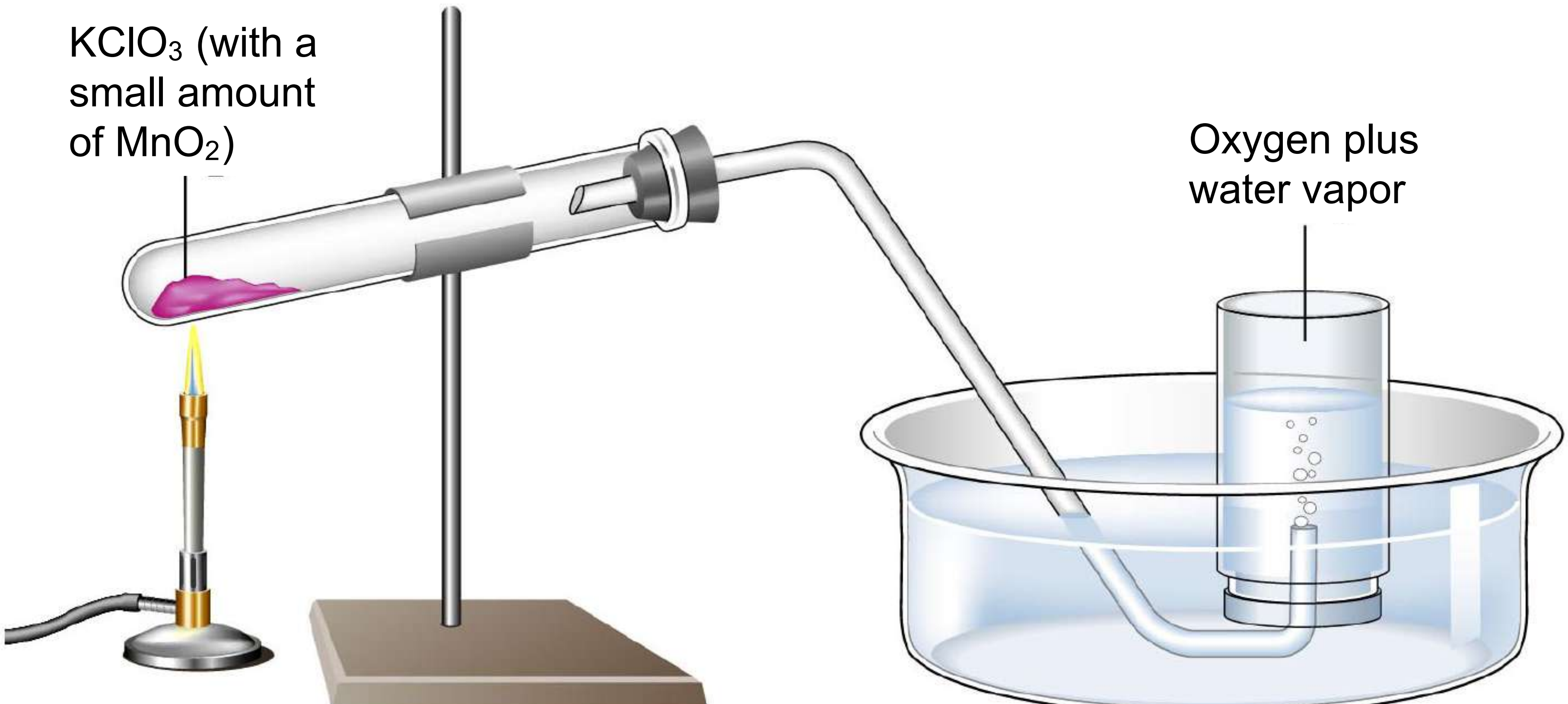
$$\chi_2 = \frac{n_2}{n_{total}} = \frac{P_2}{P_{total}}$$

The mole fraction of nitrogen in the air is 0.7808. Calculate the partial pressure of N_2 in air when the atmospheric pressure is 760. torr.

$$\chi_{N_2} \times P_{total} = P_{N_2}$$

0.7808 X 760. torr = 593 torr

The production of oxygen by thermal decomposition



Zumdahl, Zumdahl, DeCoste, World of Chemistry 2002, page 423