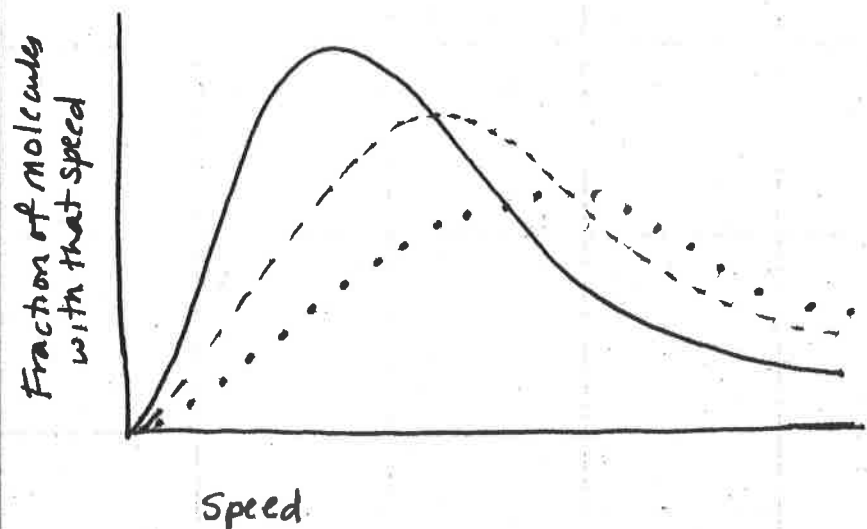


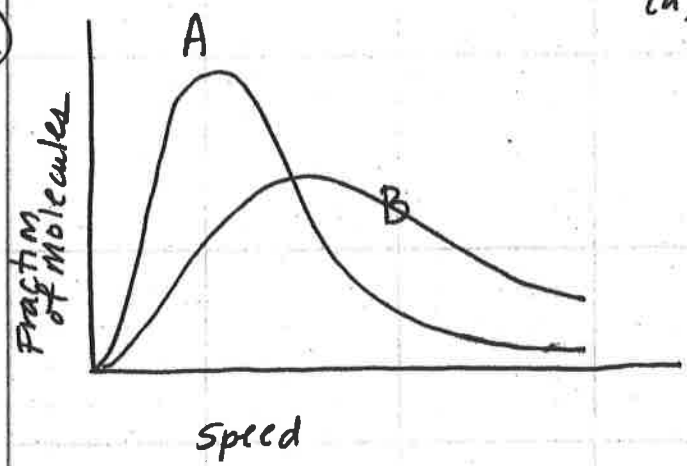
8. Sketch distribution of speeds (Fraction of molecules vs speed) for Kr @ -50°C , Kr @ 0°C , Ar @ 0°C



— Kr @ -50°C
 - - - Kr @ 0°C
 Ar @ 0°C

higher T \rightarrow higher avg speed
 lower molar mass \rightarrow higher avg speed.

9.



(a) if the ~~gas~~ curves are for He and O_2 at same temp, A = O_2 , B = He.

Helium has a lower mass, so higher average speed.

(b) if curves are for same gas at two different temperatures, A = lower Temp, B = higher Temp.

higher Temperature \rightarrow higher avg speed.

(17) Density of $H_2O = 1.00 \text{ g/mL}$
 Density of $Hg = 13.6 \text{ g/mL}$

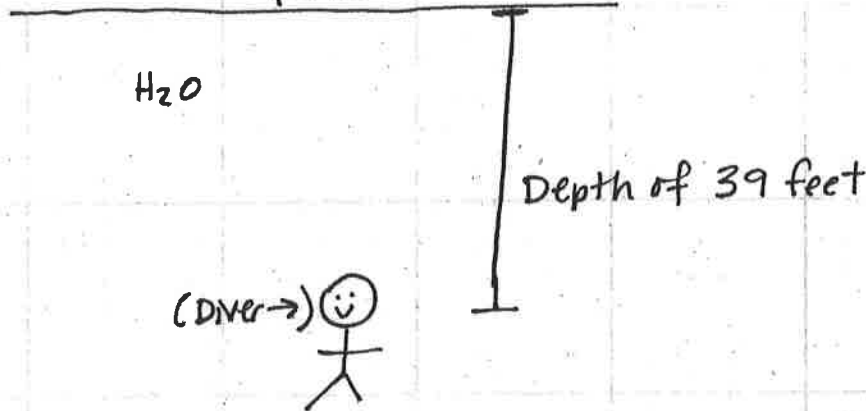
a) How tall would a column of H_2O have to be to exert a pressure equivalent to 760 mmHg ?

Since H_2O is $\frac{1}{13.6}$ th as dense as Hg , the column of H_2O would have to be 13.6 times taller than the Hg column to exert the same pressure.

$$(760. \text{ mmHg}) \left(\frac{13.6 \text{ mmHg}}{1.00 \text{ mmHg}} \right) = 10336 \text{ mmHg}$$

or 10.3 meters
 (or ≈ 34 feet!)

(b) air (atmospheric pressure = 0.97 atm)



Total pressure on diver = air pressure + pressure due to depth of H_2O

$$= 0.97 \text{ atm} + (39 \text{ ft } H_2O) \left(\frac{1.00 \text{ ft } Hg}{13.6 \text{ ft } H_2O} \right) \left(\frac{12 \text{ in}}{\text{ft}} \right) \left(\frac{25.4 \text{ mm}}{\text{inch}} \right) \left(\frac{1 \text{ atm}}{760. \text{ mmHg}} \right)$$

$$= 0.97 \text{ atm} + 1.15 \text{ atm} = 2.12 \text{ atm} \rightarrow \boxed{2.1 \text{ atm}}$$

↑
 you could also find the pressure due to H_2O by using the formula $P = D \cdot g \cdot h$

↑
 Density of the fluid

$$P = \left(\frac{1000 \text{ Kg}}{\text{m}^3} \right) \left(\frac{9.81 \text{ m}}{\text{s}^2} \right) (39 \text{ ft}) \left(\frac{.3048 \text{ m}}{\text{ft}} \right) = 116613 \frac{\text{Kg}}{\text{m}^2 \text{s}^2} \text{ or Pa}$$

$(116613 \text{ Pa}) \left(\frac{1 \text{ atm}}{101325 \text{ Pa}} \right) = 1.15 \text{ atm} \dots$ now add on the air pressure.

21. atmospheric pressure up Mt Everest (29028 ft) = 265 torr.
convert...

(a) $(265 \text{ torr}) \left(\frac{1 \text{ atm}}{760. \text{ torr}} \right) = 0.34868 \text{ atm} \rightarrow \boxed{0.349 \text{ atm}}$

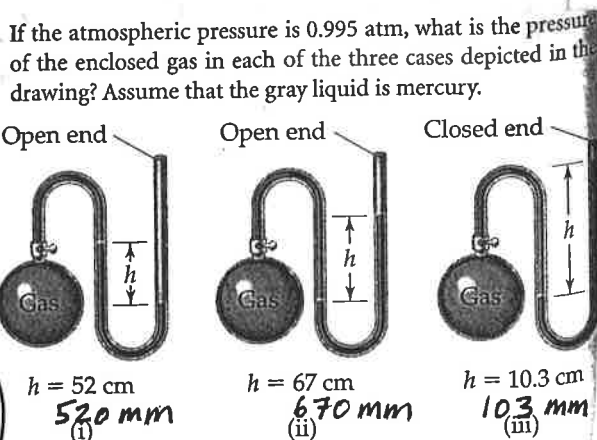
(b) $(265 \text{ torr}) \left(\frac{1 \text{ mmHg}}{1 \text{ torr}} \right) = \boxed{265 \text{ torr}}$

(c) $(0.34868 \text{ atm}) \left(\frac{101325 \text{ Pa}}{1 \text{ atm}} \right) = 35330.4 \text{ Pa} \rightarrow \boxed{35300 \text{ Pa}}$

(d) $(35330.4 \text{ Pa}) \left(\frac{1 \text{ bar}}{10^5 \text{ Pa}} \right) = 0.353 \text{ bar}$

(e) $(0.34868 \text{ atm}) \left(\frac{14.70 \text{ psi}}{1 \text{ atm}} \right) = 5.13 \text{ psi}$ (pounds per square inch)

25. In the first manometer, the gas sample ~~was~~ is at a lower pressure than the atmosphere; the atmosphere is pushing the Hg further through the tube.



(i) $P_{\text{gas}} = 0.995 \text{ atm} - (520 \text{ mmHg}) \left(\frac{1 \text{ atm}}{760. \text{ mmHg}} \right)$
 $= 0.995 \text{ atm} - 0.6842 \text{ atm}$
 $= \boxed{0.31 \text{ atm}}$

ii) The second gas has more pressure than the atmosphere; the gas is pushing the Hg further through the tube.

$P_{\text{gas}} = 0.995 \text{ atm} + (670 \text{ mmHg}) \left(\frac{1 \text{ atm}}{760. \text{ mmHg}} \right) = \boxed{1.88 \text{ atm}}$

iii) The third manometer is a closed manometer, so there is \approx no pressure above the tube on the right (\approx vacuum), so the gas has 103 mmHg more pressure than the vacuum.

$P_{\text{gas}} = 103 \text{ mmHg}$ or $\boxed{0.136 \text{ atm}}$

27) gas in a cylinder w/ piston... how will pressure be affected if...

(a) Decrease volume to $\frac{1}{4}$ original volume while keeping temp constant:
The pressure will increase by a factor of 4.

$$\left(P_1 V_1 = P_2 V_2 \quad P_2 = \frac{P_1 V_1}{V_2} = P_1 \left(\frac{V_1}{V_2} \right) = P_1 \left(\frac{V_1}{\frac{1}{4} V_1} \right) = 4 P_1 \right)$$

(b) Reduce Kelvin Temp to $\frac{1}{2}$ original value but keep V constant
Pressure will decrease by a factor of 2 (be half its original value).

$$\left(\frac{P_1}{T_1} = \frac{P_2}{T_2} \quad P_2 = P_1 \left(\frac{T_2}{T_1} \right) = P_1 \left(\frac{\frac{1}{2} T_1}{T_1} \right) = P_1 \left(\frac{1}{2} \right) \right)$$

(c) Reduce amount (molecules?) of gas to $\frac{1}{4}$ original number
while keeping V, T constant.

Pressure will decrease to $\frac{1}{4}$ its original value. (or "decrease by a factor of 4")

$\left(P = \frac{nRT}{V} \right)$ since pressure is proportional to moles of gas,
reducing moles to $\frac{1}{4}$ original value will
reduce P to $\frac{1}{4}$ of original value.)

28) Fixed Quantity of gas @ 21°C , 752 torr, 5.12 L.

(a) change P to 1.88 atm, find new V if T = const.

$$P_1 V_1 = P_2 V_2 \quad V_2 = \frac{P_1 V_1}{P_2} = \frac{(752 \text{ torr})(5.12 \text{ L})}{(1.88 \text{ atm}) \left(\frac{760 \text{ torr}}{\text{atm}} \right)} = \boxed{2.69 \text{ L}}$$

(b) Increase T to 175°C , find new V if P = constant

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad V_2 = V_1 \left(\frac{T_2}{T_1} \right) = 5.12 \text{ L} \left(\frac{(175 + 273) \text{ K}}{(21 + 273) \text{ K}} \right) = \boxed{7.80 \text{ L}}$$

29) b) 1.0 L flask of Neon vs 1.5 L flask of Xenon (at same T, P)

Avogadro's hypothesis says that gas volume is proportional to the number of gas molecules at a given temperature and pressure.

So, there are 1.5 times as many Xenon atoms/molecules in the Xenon flask as there are Neon atoms/molecules in the Ne flask.

32 (a) STP (Standard Temperature and Pressure):

$$T = 0^{\circ}\text{C} \text{ or } 273.15 \text{ K}$$

$$P = 1 \text{ atm or } 760. \text{ mmHg}$$

(b) The molar volume of an ideal gas is 22.4 L at STP.
(volume of one mole)

could find this with $PV = nRT$

$$V = \frac{nRT}{P} = \frac{(1.00 \text{ mole})(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(273.15 \text{ K})}{(1.00 \text{ atm})} = \boxed{22.4 \text{ L}}$$

(c) Find molar volume at 25°C , 1 atm

$$V = \frac{(1.00 \text{ mole})(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(298.15 \text{ K})}{(1.00 \text{ atm})} = \boxed{24.5 \text{ L}}$$

37. goodyear blimp: 175,000 ft³ of He @ 23°C , 1.0 atm.
Calculate mass of He in there.

$$n = \frac{PV}{RT} = \frac{(1.0 \text{ atm})(175000 \text{ ft}^3) \left(\frac{12 \text{ in}}{\text{ft}}\right)^3 \left(\frac{2.54 \text{ cm}}{\text{in}}\right)^3 \left(\frac{1 \text{ L}}{1000 \text{ cm}^3}\right)}{(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(296 \text{ K})}$$

$$n = 203914 \text{ moles He}$$

$$(203914 \text{ moles}) \left(\frac{4.0026 \text{ g}}{\text{mole}}\right) = 816188 \text{ grams}$$

$$\boxed{820000 \text{ g}} \text{ or } \boxed{820 \text{ kg}}$$

- (39) (a) deep breath of air: $V = 2.25 \text{ L}$, 37°C , 735 torr .
how many molecules?

$$n = \frac{PV}{RT} = \frac{\left(\frac{735}{760} \text{ atm}\right)(2.25 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(310. \text{K})} = 0.085497 \text{ moles}$$

$$(0.085497 \text{ moles}) \left(\frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mole}} \right) = \boxed{5.15 \times 10^{22} \text{ molecules}}$$

- (b) Whale Lung capacity = $5.0 \times 10^3 \text{ L}$ (!)
mass air in lungs @ 0.0°C , 1.00 atm . use 28.98 g/mole for molar mass.
This is at STP so don't have to use $PV=nRT$...

$$(5.0 \times 10^3 \text{ L}) \left(\frac{1 \text{ mole}}{22.4 \text{ L}} \right) \left(\frac{28.98 \text{ g}}{\text{mole}} \right) = 6469 \text{ grams} \downarrow$$

$$\boxed{6500 \text{ g}} \text{ or } \boxed{6.5 \text{ kg}}$$

- (c) Air is $\approx 78\% \text{ N}_2$, $21\% \text{ O}_2$, $1\% \text{ Ar}$, (by moles)

so molar mass would be $\approx (0.78)(28.0134 \frac{\text{g}}{\text{mole}})$

$$+ (0.21)(31.9988 \frac{\text{g}}{\text{mole}})$$

$$+ (0.01)(39.948 \frac{\text{g}}{\text{mole}})$$

$$= 28.97 \approx \boxed{29 \text{ g/mole}}$$

- (46) (a) $V_{\text{O}_2 \text{ max}} = \frac{47.5 \text{ mL O}_2}{\text{kg body mass} \cdot \text{minute}}$ for a 185 lb man.

Volume of O_2 consumed in one hr (at $V_{\text{O}_2 \text{ max}}$ rate!)

$$\left(\frac{47.5 \text{ mL O}_2}{\text{kg} \cdot \text{min}} \right) (185 \text{ lb}) \left(\frac{1 \text{ kg}}{2.2046 \text{ lb}} \right) \left(\frac{60 \text{ min}}{\text{hr}} \right) = 239159 \text{ mL}$$

$$2.39 \times 10^5 \text{ mL}$$

- (b) Convert to grams O_2 and pounds O_2 ,
if at 1 atm and 25°C

or 239 Liters

$$n = \frac{PV}{RT} = \frac{(1 \text{ atm})(239.159 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298 \text{ K})} = 9.7752 \text{ moles O}_2$$

$$(9.7752 \text{ moles}) \left(\frac{31.9988 \text{ g}}{\text{mole}} \right) = 312.795 \text{ g} \rightarrow \boxed{313 \text{ grams O}_2}$$

$$(312.795 \text{ g}) \left(\frac{1 \text{ lb}}{453.59 \text{ g}} \right) = \boxed{0.690 \text{ pounds of O}_2}$$

(50)	SO ₂ 64 amu #2	HBr 81 amu most dense #1	CO ₂ 44 amu least dense #3
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At a given temperature and pressure, gas density is proportional to molar mass.

$$D = \frac{m}{V} = \frac{\text{molar mass} \cdot n}{V} = \frac{\text{mm} \cdot n}{\frac{nRT}{P}}$$

$$D = \frac{\text{mm} \cdot P}{RT}$$

(#52) (b) is correct

(53) a) Density of NO₂ @ 0.970 atm, 35°C

Volume of 1 mole : $V = \frac{nRT}{P} = \frac{(1.00 \text{ mole})(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(308 \text{ K})}{(0.970 \text{ atm})}$

$$V = 26.069 \text{ L}$$

$$D = \frac{m}{V} = \frac{46.0055 \text{ g/mole}}{26.069 \text{ L/mole}} = \boxed{1.76 \text{ g/L}}$$

(b) 2.50 g of gas occupies 0.875 L @ 685 torr, 35°C. find molar mass

$$n = \frac{PV}{RT} = \frac{(\frac{685}{760} \text{ atm})(0.875 \text{ L})}{(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(308 \text{ K})} = 0.031188 \text{ moles}$$

$$\frac{2.50 \text{ g}}{0.031188 \text{ moles}} = \boxed{80.2 \text{ g/mole}}$$

(54) a) Density of SF₆ @ 707 torr, 21°C

molar volume: $V = \frac{nRT}{P} = \frac{(1 \text{ mole})(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(294 \text{ K})}{(\frac{707}{760}) \text{ atm}} = 25.947 \text{ Liters}$

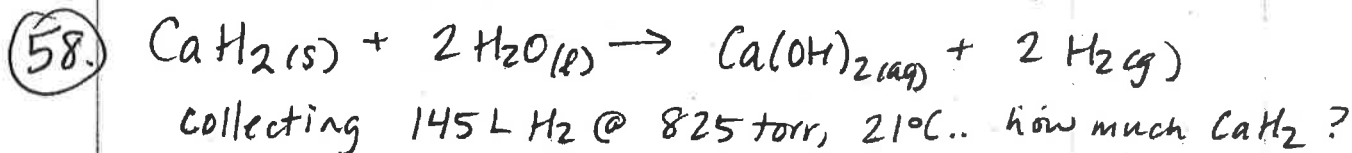
$$D = \frac{m}{V} = \frac{146.0554 \text{ g/mole}}{25.947 \text{ L/mole}}$$

$$= \boxed{5.63 \text{ g/L}}$$

(b) if D = 7.135 g/L @ 12°C, 743 torr, find molar mass.

Volume of 1 mole = $\frac{(1 \text{ mole})(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(285 \text{ K})}{(\frac{743}{760} \text{ atm})} = 23.934 \text{ L}$

$$\left(\frac{7.135 \text{ g}}{\text{L}}\right) \left(\frac{23.934 \text{ L}}{\text{mole}}\right) = \boxed{171 \text{ g/mole}}$$



$$n_{\text{H}_2} = \frac{PV}{RT} = \frac{\left(\frac{825}{760} \text{ atm}\right)(145 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(294 \text{ K})} = 6.5211 \text{ moles H}_2$$

$$(6.5211 \text{ moles H}_2) \left(\frac{1 \text{ mole CaH}_2}{2 \text{ mole H}_2}\right) \left(\frac{42.0958 \text{ g}}{\text{mole}}\right) = \boxed{137 \text{ g CaH}_2}$$



(a) 24.5g

? Volume
at 37°C
0.970 atm

$$(24.5 \text{ g C}_6\text{H}_{12}\text{O}_6) \left(\frac{1 \text{ mole}}{180.15588 \text{ g}}\right) \left(\frac{6 \text{ mole CO}_2}{1 \text{ mole C}_6\text{H}_{12}\text{O}_6}\right) = 0.815960 \text{ moles CO}_2$$

$$V = \frac{nRT}{P} = \frac{(0.815960 \text{ mole}) \left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right) (310. \text{ K})}{(0.970 \text{ atm})} = \boxed{21.4 \text{ L CO}_2}$$

(b) Volume O_2 needed @ 1.00 atm, 298 K to oxidize 50.0 g glucose

$$(50.0 \text{ g C}_6\text{H}_{12}\text{O}_6) \left(\frac{1 \text{ mole}}{180.15588 \text{ g}}\right) \left(\frac{6 \text{ mole O}_2}{1 \text{ mole C}_6\text{H}_{12}\text{O}_6}\right) = 1.6652 \text{ mole O}_2$$

$$V = \frac{nRT}{P} = \frac{(1.6652 \text{ mol}) \left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right) (298 \text{ K})}{(1.00 \text{ atm})} = 40.7 \text{ L O}_2$$

(since air is only 21% O_2 , you'd need 194 Liters of air!)

(50.0 g glucose \approx 100 food calories)



$$P_{\text{H}_2} = P_{\text{barometric}} - V P_{\text{H}_2\text{O}} @ 24^\circ\text{C}$$

@ 24°C, 738 torr baromet.

$$P_{\text{H}_2} = 738 \text{ torr} - 22.38 \text{ torr} \leftarrow \text{from Appendix B}$$

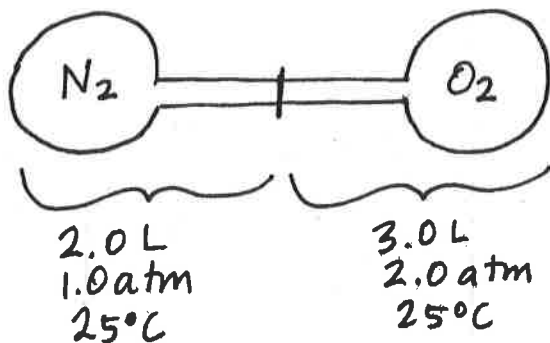
$$P_{\text{H}_2} = 715.62 \text{ torr}$$

$$n_{\text{H}_2} = \frac{P_{\text{H}_2} V}{RT} = \frac{\left(\frac{715.62}{760} \text{ atm}\right) (0.159 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right) (297 \text{ K})} = 0.00613997 \text{ moles H}_2$$

$$(0.00613997 \text{ mole H}_2) \left(\frac{1 \text{ mole Zn}}{1 \text{ mole H}_2}\right) \left(\frac{65.39 \text{ g}}{\text{mole}}\right) = \boxed{0.401 \text{ g Zn}}$$

(.40149)

63



(a) when valve is opened, the N_2 volume increases from 2.0 L to 5.0 L (it will spread out into the entire container)

so P_{N_2} will decrease: $P_1 V_1 = P_2 V_2$

$$P_2 = \frac{P_1 V_1}{V_2} = \frac{(1.0 \text{ atm})(2.0 \text{ L})}{(5.0 \text{ L})} = \boxed{0.40 \text{ atm}} \quad N_2$$

(b) when valve is opened, the O_2 will be able to occupy 5.0 L instead of 3.0.

$$P_2 = \frac{P_1 V_1}{V_2} = \frac{(2.0 \text{ atm})(3.0 \text{ L})}{(5.0 \text{ L})} = \boxed{1.2 \text{ atm}} \quad O_2$$

$$(c) P_{\text{TOTAL}} = P_{N_2} + P_{O_2} = 0.40 \text{ atm} + 1.2 \text{ atm} = \boxed{1.6 \text{ atm}}$$

85

(a) given:

Ne	HBr	SO ₂	NF ₃	CO
20	80	64	71	28
amu	amu	amu	amu	amu

Since they are all at the same temperature (25°C) they should have the same avg kinetic energy. (KE)
 since $KE = \frac{1}{2} m v^2$, the molecules with lower mass (m) will have higher speed (v) in order to have same KE.

So, in order from slowest to fastest:

HBr, NF₃, SO₂, CO, Ne

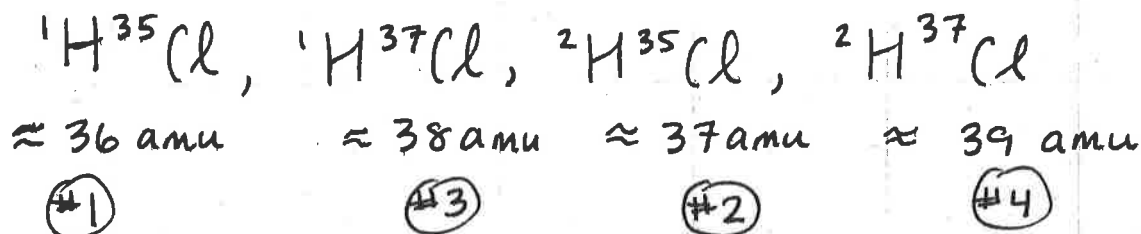
(b) $V_{\text{rms}} = \sqrt{\frac{3RT}{M}}$ NF₃: $M = 71.0019 \text{ g/mole}$

$$= \left(\frac{3(8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}})(298 \text{ K})}{.0710019 \text{ kg/mole}} \right)^{1/2} = \left(104683 \frac{\text{m}^2}{\text{s}^2} \right)^{1/2}$$

$$= 323.55 \text{ m/s} \rightarrow \boxed{324 \text{ m/s}}$$

89) Hydrogen isotopes: ${}^1_1\text{H}$ and ${}^3_1\text{H}$
 Chlorine isotopes: ${}^{35}_{17}\text{Cl}$ and ${}^{37}_{17}\text{Cl}$

Consider relative rates of effusion of



↑
 fastest / highest rate of effusion (escape through a tiny hole)
 since lowest molar mass will correspond to higher speed,
 and therefore more frequent "collisions" with the hole.

91) Arsenic (III) sulfide is a solid @ room temp,
 but sublimates readily.

Must have empirical formula of As_2S_3 based on charges,
 but has large ore gas molecules?

The vapor's effusion rate is 0.28 the rate of effusion for Ar.

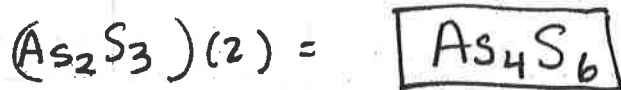
$$\frac{r_{\text{gas}}}{r_{\text{Ar}}} = \frac{v_{\text{gas}}}{v_{\text{Ar}}} = \left(\frac{M_{\text{Ar}}}{M_{\text{gas}}} \right)^{1/2} = 0.28$$

$$\frac{M_{\text{Ar}}}{M_{\text{gas}}} = (0.28)^2 = 0.0784$$

$$M_{\text{gas}} = \frac{M_{\text{Ar}}}{0.0784} = \frac{39.948}{0.0784} = 509.5 \text{ amu} \rightarrow 510 \text{ amu}$$

As_2S_3 : empirical mass = 246 amu

$2(246) = 492 \text{ amu}$.. as close as we'll get to 510 amu



92

a gas w/ unknown molar mass : 105s for 1.0 L to effuse at same conditions, O₂ needs 31s for 1.0 L to effuse.

effusion rate is inversely proportional to effusion time,

$$\text{so } \frac{r_{O_2}}{r_x} = \frac{t_x}{t_{O_2}} = \frac{105s}{31s} = 3.3871$$

$$\frac{r_{O_2}}{r_x} = \left(\frac{M_x}{M_{O_2}} \right)^{1/2} = 3.3871$$

$$\frac{M_x}{M_{O_2}} = (3.3871)^2 = 11.472$$

$$M_x = M_{O_2}(11.472) = 31.9988(11.472) = 367.9 \text{ g/mole}$$

370 g/mole

93

(a) gases deviate most from ideal behavior @ low Temperatures and high Pressures.

(b) ^{how} gases deviate from ideal behavior:

(1) An "ideal" gas would not have any intermolecular attractive forces (and therefore molecules would never condense to a liquid). A real gas has attractive forces. These are more "noticeable" at low T and high P. At low T, the molecules have less KE (Kinetic energy), so colliding/close molecules have more time to form attractive forces. At high P, molecules are closer to each other on avg, so again, have more opportunity to have attractive forces.

(2) In an ideal gas, the volume taken up by the atoms/molecules themselves is negligible compared to the overall sample/container volume. In a real gas, the molecules do themselves have volume. this becomes more noticeable at high P and low T... basically it is more noticeable at high $\frac{P}{V}$ values, since this involves more molecules in a given V....

96) Van der Waals equation: $(p + \frac{an^2}{V^2})(V - nb) = nRT$

"a" accounts for intermolecular attractive forces (IMF).

$$p = \frac{nRT}{V - nb} - \frac{an^2}{V^2}$$

← for larger value of a, Pressure will be reduced, as molecules stick to each other more (stronger IMF) and hit the walls of the container with less force.

"b" accounts for the volume taken up by the molecules themselves. The larger the "b", the more space the molecules themselves occupy; this reduces the "free" volume available for molecules, and increases pressure.

109) cylinder of a car engine: $V = 524 \text{ cm}^3$

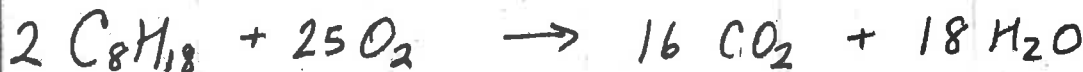
(a) air at 74°C , 0.980 atm (air is 20.95% O_2 by moles).
Find moles O_2

$$n_{\text{air}} = \frac{P_{\text{air}} V}{RT} = \frac{(0.980 \text{ atm})(0.524 \text{ L})}{(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(347 \text{ K})} = 0.018025 \text{ moles air}$$

but air is only 20.95% O_2 , so $\rightarrow (0.018025)(0.2095) = 0.003776 \text{ moles } \text{O}_2$

$0.00378 \text{ moles } \text{O}_2$

(b) what mass of C_8H_{18} (in grams) could be combusted w/ this amount of O_2 ?



$$(0.003776 \text{ mole } \text{O}_2) \left(\frac{2 \text{ mole } \text{C}_8\text{H}_{18}}{25 \text{ mole } \text{O}_2} \right) \left(\frac{114.2278 \text{ g}}{1 \text{ mole}} \right) = 0.0345 \text{ g } \text{C}_8\text{H}_{18}$$

115) gas sample @ -33°C .

What Temp is needed to increase v_{rms} by factor of 2?

$$KE = \frac{1}{2} mV^2$$

if we double the speed, the KE will quadruple, which means we need to quadruple the Kelvin temp, since KE is proportional to the absolute temp.

$$(-33^\circ\text{C} + 273) \times 4 = 960. \text{ Kelvin}$$

$$960. \text{ K} - 273 = 687^\circ\text{C}$$

(or use $v = \sqrt{\frac{3RT}{M}}$ to figure this out)

(117)

(a) compressing a gas to a smaller volume @ constant T will cause the IMF to have a larger effect on the properties of a gas, because molecules will be closer together, and have more opportunity to attract e/o.

(b) Increasing Temp while keeping V constant will decrease the effect of the IMF. When molecules collide, they will be close enough to experience IMF for shorter periods of time / will be more able to "escape" the IMF.

(12) "Cyclopropane" 85.7% C, 14.3% H by mass.

1.56 g C_xH_y occupies 1.00 L @ 0.984 atm, 50.0°C

(a) Molecular Formula = ?

$$(85.7g C) \left(\frac{1 \text{ mole}}{12.011g} \right) = 7.1351 \text{ moles C}$$

$$\frac{14.1879}{7.1351} = 1.988 \approx 2$$

$$(14.3g H) \left(\frac{1 \text{ mole}}{1.0079g} \right) = 14.1879 \text{ moles H}$$

CH_2 = empirical formula
(14 amu)

$$n = \frac{PV}{RT} = \frac{(0.984 \text{ atm})(1.00 \text{ L})}{(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(323 \text{ K})} = 0.037106 \text{ moles}$$

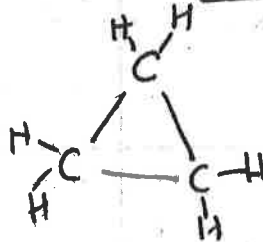
$$\frac{1.56 \text{ g}}{0.037106 \text{ mole}} = 42.0 \text{ g/mole.}$$

$$14(3) = 42$$

$$(CH_2)(3) = \boxed{C_3H_6}$$

Remember this?

Cyclopropane
 ↖ ring ↑ ↗ single bonds
 3 carbons



(c) How would cyclopropane's effusion rate compare to methane's? (methane = CH_4)

cyclopropane will have a slower effusion rate than CH_4 .
 C_3H_6 has a larger molar mass and therefore a slower speed and effusion rate, than CH_4 .

Chapter 10 #131

10.131 a. Compare the effusion rates of neon and radon, if the gases are at the same temperature. (Which gas has the higher effusion rate, and by what factor?)

$$\frac{r_{\text{Ne}}}{r_{\text{Rn}}} = \sqrt{\frac{M_{\text{Rn}}}{M_{\text{Ne}}}} = \sqrt{\frac{222 \text{ amu}}{20.179 \text{ amu}}} = 3.3169 \rightarrow 3.32$$

$$\boxed{\frac{r_{\text{Ne}}}{r_{\text{Rn}}} = 3.32}$$

(Neon's effusion rate is 3.32 times faster than Radon's.)

b. A compound with a super creative name ("Compound X") is tested, and its effusion rate is compared to that of Argon (Ar). When compound X and Argon are at the same temperature (and both in gas phase), Argon's rate of effusion is 1.73 times higher than that of compound X. Use this info to find the molar mass of compound X.

$$\frac{r_{\text{Argon}}}{r_{\text{X}}} = \sqrt{\frac{M_{\text{X}}}{M_{\text{Argon}}}} = 1.73$$

$$\sqrt{\frac{M_{\text{X}}}{39.948}} = 1.73$$

$$\frac{M_{\text{X}}}{39.948} = (1.73)^2$$

$$M_{\text{X}} = 119.56 \rightarrow \boxed{120. \frac{\text{g}}{\text{mole}}}$$

c. If compound X has the empirical formula CH_2O , what is the molecular formula of this compound?

$$\text{empirical mass of } \text{CH}_2\text{O} \approx 1(12) + 2(1) + 1(16) = 30. \text{ amu or g/mole}$$

$$\frac{120.}{30.} = 4$$

