CHEM 101A – Chapter 5 Gas Problems

$T_K = T_C + 273$

1. Oxygen gas occupies 573 mL (V_1) at 50.°C (T_1). If the temperature is reduced to -78°C (T_2), what is the volume of gas, assuming pressure and moles are held constant?

$$\frac{573\,mL}{323\,K} = \frac{V_2}{195\,K}$$

$$V_2 = 345.9287$$

$$\boxed{346\,mL}$$

2. A sample of an ideal gas is placed in a 20.0-L vessel at STP. The volume is reduced to 500. mL whilst the temperature is increased to 150.°C. Find the pressure.

$P_1 = 1\,atm$ $P_2 = ?$

$V_1 = 20.0\,L$ $V_2 = 0.500\,L$

$T_1 = 273\,K$ $T_2 = 423\,K$

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

$$P_2 = \frac{T_2 P_1 V_1}{V_2 T_1} = \frac{(423\,K)(1\,atm)(20.0\,L)}{(0.500\,L)(273\,K)}$$

$$P_2 = 61.9780$$

$$\boxed{P_2 = 62.0\,atm}$$

3. An unknown elemental gas has a pressure of 722 mm Hg at 100.°C and has a density of 2.60 g/L. What is the identity of the gas? *Hint: if it is a diatomic element, its molar mass is twice its atomic mass!*

$722\ mmHg \times \frac{1\ atm}{760\ mmHg} = 0.950\ atm$

$MP = dRT$

$M = \frac{dRT}{P} = \frac{(2.60\ \frac{g}{L})(.08206\ \frac{L \cdot atm}{K \cdot mol})(373\ K)}{0.950\ atm}$

$M = 83.8\ g/mol = Kr$

4. Determine the densities of the noble gases at STP.

5. Concentrated H_2O_2 solutions are explosively decomposed by transition metal catalysts.

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

?L

What volume of dry O_2 collected @27°C and 746 torr could be generated by decomposition of 125g H_2O_2(50.0% by mass)?

(Sample mass)(% mass) = mass pure

$$(125g\ H_2O_2\ 50\%)(.500) = 62.5g\ H_2O_2$$

$$62.5g\ H_2O_2 \times \frac{1\ mol\ H_2O_2}{34.016\ g\ H_2O_2} \times \frac{1\ mol\ O_2}{2\ mol\ H_2O_2} = 0.919\ mol\ O_2 = n$$

$$PV = nRT$$

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

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

$$P = 746\ torr \times \frac{1\ atm}{760\ torr} = 0.9816\ atm$$

$$T = 273 + 27 = 300.\ K$$

$$= \frac{(.919\ mol\ O_2)(.08206\ \frac{L \cdot atm}{K \cdot mol})(300.\ K)}{0.9816\ atm}$$

$$= 23.0\ L\ O_2$$

6. Urea (H2NCONH2) can be synthesized from ammonia and carbon dioxide:

$$2NH_3(g) + CO_2(g) \rightarrow H_2NCONH_2(s) + H_2O(g)$$

(2)

At 223°C, ammonia gas at 90.atm flows into a reactor at a rate of 500.L/min whilst carbon dioxide gas at 45 atm flows into a reactor at a rate of 600.L/min. What mass urea is produced per minute at 100% yield?

L.R

$$n_{NH_3} = \frac{PV}{RT} = \frac{(90.\text{atm})(500.\text{L})}{(.08206\ \frac{\text{L}\cdot\text{atm}}{\text{K}\cdot\text{mol}})(496\ \text{K})} = 1105.6 \text{ mol } NH_3$$

$$n_{CO_2} = \frac{(45)(600)}{(.08206)(496)} = 663.4 \text{ mol } CO_2$$

X.S.

$$1105.6 \text{ mol } NH_3 \times \frac{1 \text{ mol Urea}}{2 \text{ mol } NH_3} \times \frac{60.06 \text{ g Urea}}{1 \text{ mol Urea}} =$$

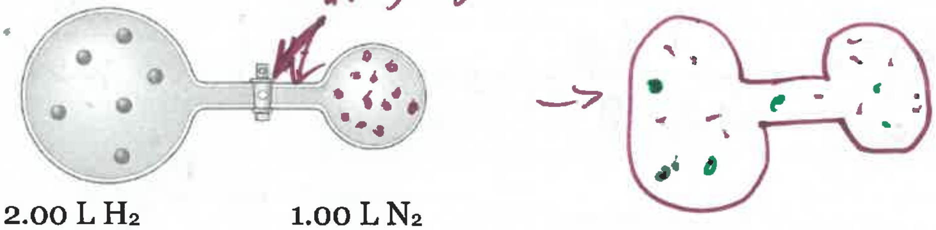
$$= 33,202 \text{ g Urea}$$

$$= 33,000 \text{ g Urea}$$

$$= 33 \text{ kg Urea/min}$$

7. Consider the 2-bulbed flask below. What are the final partial pressures (in torr) of H_2 and N_2 after the stopcock is opened? *Hint: try Boyle's Law*

negligible



2.00 L H_2	1.00 L N_2
475 torr	2.00 atm

→ 1520 torr

$P_1V_1 = P_2V_2$

H_2: P_1 = 475 torr, V_1 = 2.00 L → V_2 = 3.00 L, P_2 = ? torr ↓

$P_{H_2} = 317$ torr

N_2: 1.00 L → 3.00 L; 1520 torr → ? torr ↓

$P_{N_2} = 507$ torr

P_{tot} = 317 torr + 507 torr
= 824 torr

8. You have two balloons—one filled with hydrogen and one with xenon gas, both at 298K.
   a. What is the root mean square velocity of the hydrogen molecules? (H_2)

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$

UNITS

$$u_{rms} = \sqrt{\frac{\frac{kg \cdot m^2/s^2}{K \cdot mol} \cdot \frac{K}{1}}{\frac{kg}{mol}}}$$

$$\frac{2.016 g H_2}{mol} \times \frac{1 kg}{1000 g} = .002016 \ kg/mol$$

$$u_{rms} = \sqrt{\frac{3 \cdot 8.314 \cdot 298}{.002016}} = \frac{m}{s} = \sqrt{3,686,863}$$

$$u_{rms} = 1920 \ m/s$$

~4000 mi/h

Same $(KE)_{avg}$

   b. What is the root mean square velocity of the xenon atoms?

$$u_{rms} = \sqrt{\frac{3 \cdot 8.314 \cdot 298}{0.1313}} = 238 \ m/s$$

SF6 H2O
He

9. Calculate the average kinetic energy of one molecule of CO_2 at 150°C.

$$(KE)_{avg} = \frac{3}{2}RT$$

$$= \frac{3}{2}\left(8.314\ \frac{J}{K\cdot mol}\right)(423\ K)$$

$$= 5280\ \frac{J}{mol} \times \frac{1\ mol}{6.022\times10^{23}\ molecule}$$

$$= 8.76\times10^{-21}\ J/molecule$$

10. Calculate the pressure exerted by 1.500 moles of xenon in a 2.000-L flask at a) 200.0K and b) 1000. K, using the ideal gas equation and the Van der Waals equation.

IDEAL

T= 200.0K $P = \frac{(1.500)(.082057)(200.0)}{2.000} =$

T= 1000. K $P = \frac{(1.500)(.082057)(1000.)}{2.000} =$

VDW

T=200.0K

$$P = \frac{(1.500)(.082057)(200.0)}{(2.000 - (1.500)(.0511))} - \left(\frac{1.500}{2.000}\right)^2 \cdot 4.19$$

= 10.44 atm

T=1000.K

$$P = \frac{(1.500)(.082057)(1000.)}{(2.000 - 1.500 \cdot .0511)} - \left(\frac{1.500}{2.000}\right)^2 \cdot 4.19$$

= 61.64atm