Gas review key from Chem. 120

1. A 6.75 L flask contains a fixed amount of gas at 31oC and a constant pressure. If the temperature is increased to 125oC, what will the volume of the gas be?

 *As temperature increases, Volume should increase if pressure and amount are constant.*

1. A vacuum pump exhausts a heavy-walled 1.50-L round-bottomed flask to a pressure of 3.55 x 10-6torr. How many particles are present if the temperature is 273 K?

$$PV=nRT\rightarrow n=\frac{PV}{RT}=\frac{\left(3.55 ×10^{-6}torr\right)\left(1.50 L\right) mol K}{\left(62.4 L torr\right)\left(273 K\right)}$$

$$=3.12×10^{-10}mol particles$$

$$?particles=3.12×10^{-10}mol×\frac{6.022×10^{23}particles}{1 mol}=1.88 ×10^{14} particles$$

How many molecules of xenon hexafluoride are required to react with 0.579 L of hydrogen gas at 2.46 atm and 45 °C in the unbalanced reaction show below (10 points)?

XeF6 (s) + 3 H2 (g) 🡪 Xe (s) + 6 HF (g)

V = 0.579 L

P = 2.46 atm

T = 45 ° + 273 = 318 K

$PV=nRT⟹n=\frac{PV}{RT}=\frac{\left(2.46 atm\right)(0.579 L)}{\left(0.0821 \frac{L atm}{mol H\_{2} K}\right)(318 K)}×\frac{1 mol XeF\_{6}}{3 mol H\_{2}}×\frac{6.022 × 10^{23}molecule XeF\_{6}}{1 mol XeF\_{6}}=1.10 ×10^{22}molecules XeF\_{6}$

1. It is found that 250. mL of an unknown ideal gas at STP has a mass of 2.50 g. What is the molar mass of the unknown gas?

 V = 0.250 L, T = 273 K, P = 1 atm, mass = 2.50 g what is n?

 $Moles = \frac{PV}{nT }= \frac{1.00 atm ×(0.250 L) }{\left(0.0821 \frac{L atm}{mol K}\right)×273 K}=0.0112 mol gas $

$$MM= \frac{2.50 g gas}{ 0.0112 mol gas}= \frac{224 g}{mol}$$

 Or

 $0.250 L × \frac{1 mol gas}{22.4 L}= 0.0112 mol gas$

 $MM= \frac{2.50 g gas}{ 0.0112 mol gas}= \frac{224 g}{mol}$

1. A mixture of hydrogen, H2, nitrogen, N2, and Argon, Ar, gases is present in a steel cylinder. The total pressure within the cylinder is 675 mm Hg and partial pressures of nitrogen and argon gases are, respectively, 354 mm Hg and 235 mmHg. If carbon dioxide, CO2, gas is added to the mixture at constant temperature until the total pressure reaches 842 mm Hg, what is the partial pressure, in millimeters of Hg, of the following (8 points)?
	1. CO­2

$$P\_{Total 1}=P\_{H\_{2}}+P\_{N\_{2}}+P\_{Ar}=675 mm Hg$$

$$P\_{Total 2}=P\_{Total 1}+P\_{CO\_{2}}⇒P\_{CO\_{2}}=P\_{Total 2}-P\_{Total 1}$$

$$P\_{CO\_{2}}=842 mm Hg-675 mm Hg=167 mm Hg$$

* 1. N2

$$P\_{N\_{2}}=354 mm Hg$$

* 1. Ar

$$P\_{Ar}=235 mm Hg$$

* 1. H2

$$P\_{H\_{2}}=P\_{Total 1}-P\_{N\_{2}}-P\_{Ar}=675 mm Hg-354 mm Hg-235 mm Hg=86 mm Hg$$

1. Nitrogen dioxide gas reacts with water vapor to produce oxygen and ammonia gases. Suppose that 12.8 g of nitrogen dioxide reacts with 5.00 L of water vapor at 375 °C and 725 torr (24 points).
2. Write the balanced chemical equation

4 NO2 (g) + 6 H2O (g) 🡪 7 O2 (g) + 4 NH3 (g)

1. Use an ICE table to find the limiting reagent and complete the table

4 NO2 (g) + 6 H2O (g) 🡪 7 O2 (g) + 4 NH3 (g)

I 0.278 mol 0.0897 mol 0.000 mol 0.000 mol

C -4x -6x +7x +4x

E 0.278 mol – 4x = 0.0897 mol – 6x = 0.00 +7x = +4x

 0.278 mol – 4(0.0150 mol) = 0.0897 mol = 6x 7(0.0150 mol) = 4(0.0150 mol) =

 0.278 mol – 0.0600 mol = x = 0.0150 mol 0.105 mol O2 0.0600 mol NH3

 0.218 mol NO2

$$12.8 g NO\_{2}×\frac{1 mol NO\_{2}}{46.02 g NO\_{2}}=0.278 mol NO\_{2}$$

H2O

V = 5.00 L

P = 725 torr

T = 375 °C + 273 = 648 K

$$PV=nRT ⇒n=\frac{PV}{RT}=\frac{\left(725 torr\right)(5.00 L)}{\left(62.4 \frac{L torr}{mol K}\right)(648 K)}=0.0897 mol H\_{2}O$$

1. How many grams of ammonia are produced?
	1. $mol NH\_{3}×\frac{17.04 g NH\_{3}}{1 mol NH\_{3}}=1.02 g NH\_{3}$
2. What is the percent yield if 0.984 g NH3 are collected?

$$\%yield=\frac{actual yield}{theoretical yield}×100\%=\frac{0.984 g}{1.02 g}×100\%=96.5\% NH\_{3}$$

1. How many molecules of the excess reagent remain?
	1. $mol NO\_{2}×\frac{6.022×10^{23} molecules NO\_{2}}{1 mol NO\_{2}}=1.31×10^{23} molecules NO\_{2}$

Newer material

1. What is the pressure, in atmospheres, of the gas in the following open-tube manometer? Assume that atmospheric pressure is 102 kPa. (5 points)



102 kPa (1 atm / 101 kPa) = 1.01 atm

27 mm Hg (1 atm / 760 mm Hg) = 0.0355 atm

Pressure = 1.01 – 0.0355 = 0.975 atm OR 98.5 kPa

1. An artificial atmosphere for the planet Jupiter consists of 92% hydrogen and 8% helium. 15g of this gas mixture are placed in a 50L container at –25°C. What is the partial pressure of each gas? (5 points) What is the mole fraction of hydrogen and helium in the artificial atmosphere? (5 points)

Hydrogen (H2): (15g \* 0.92)(1 mol/2.02g) = 6.83 mol

Helium (He): (15g \* 0.08)(1 mol/4.00g) = 0.3 mol

PV = nRT

P = nRT/V

P(H2) = 6.83 mol \* 0.0821 mol L/atm K \* (273-25K) / 50L = 2.78 atm

P(He) = 0.3 mol \* 0.0821 mol L/atm K \* (273-25K) / 50L = 0.122 atm

Mole fraction Hydrogen = 6.83/(6.83+0.3) = 0.958

Mole fraction Helium = 0.3/(6.83+0.3) = 0.042

1. **SHOW ALL WORK.** You are asked to select a high precision valve that will be used to smoothly deliver uranium hexafluoride (UF6) at a rate of 0.025 L/min. For safety and economic reasons, you decide to use helium (He) to test the flow rate of the new valve. Determine the correct He flow rate of this valve in L/min.

ERHe / ERUF6 = (FMUF6 / FMHe)1/2 (Graham's Law of Effusion)

ERHe / (0.025 L/min) = (352 / 4)1/2

ERHe = 0.235 L/min

1. For a spacecraft or a molecule to leave the moon, it must reach the escape velocity (speed) of the moon, which is 2.37 km/s. The average daytime temperature of the moons surface is 365 K. What is the rms speed (in m/s) of a hydrogen molecule at this temperature? How does this compare with the escape velocity?

urms = (3RT/*M*)1/2 = [3(8.3144 J/K-mol)(365 K)/(0.002016 kg/mol)]1/2  = **2,130 m/s**

**The rms velocity of a hydrogen molecule is lower than the escape velocity, but not by much. So a very large fraction of the hydrogen molecules would be moving at velocities higher than the escape velocity. Over a short period of time, any hydrogen gas on the moon will have escaped into space.**

1. A given volume of nitrogen, N2, required 68.3 s to effuse from a hole in a chamber. Under the same conditions, another gas required 85.6 s for the same volume to effuse. What is the molecular weight of this gas?

t1/t2 = (*M*1/*M*2)1/2 call unknown "gas 1" and call N2 "gas 2", then...

*M*1 = (t1/t2)2*M*2 = (85.6 s/68.3 s)2(28.014 g/mol) = **44.0 g/mol**
2. Calculate the pressure of water vapor at 120.0° C if 1.000 mol of water vapor occupies 32.50 L. Use the van der Waals equation (see Table 5.7 for data). Compare the result from the ideal gas law.  (a= 5.537 L2-atm/mol2 and b = 0.03049 L/mol)

          Now plug appropriate values into van der Waals equation, simplify, then solve for P.

          (P + (5.537)(1.000)2/(32.50 L)2)(32.50 L - (1.000)(0.03049)) = (1.000)(0.082057)(393.15 K)

          (P + 0.005242 atm)(32.50 L - 0.03049 L) = 32.26 L-atm

          (P + 0.005242 atm)(32.4695 L) = 32.26 L-atm

          (P + 0.005242 atm) = 0.993547 atm

**P = 0.9883 atm by van der Waals equation**

          P = nRT/V = (1.000 mol)(0.082057 L-atm/K-mol)(393.15 K)/32.50 L)

**= 0.9926 atm by ideal gas law**

**The value for the pressure from the van der Waals equation is about the same, but slightly lower than, the value from the ideal gas law. At low pressures, this is how it should be.**