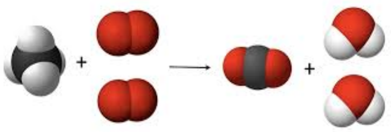
5.4 Gas Stoichiometry



$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$

Molar Volume and STP

Avogadro's Law: equal volume of gases at the same temperature and pressure contain equal numbers of particles.

1. At STP (0 degrees Celsius (273 K), and 1.0 atm pressure), what volume will one mole of gas occupy?

$$PV = nRT$$

$$V = \frac{nRT}{P} = \frac{1.00\,mol \bullet 0.0821\frac{L\bullet atm}{K\bullet mol}\bullet 273\,K}{1.00\,atm} = \boxed{22.4\,L}$$

2. Molar Volume Example

What volume will 2.16 g of krypton gas occupy at STP?

Convert mass to moles:

$$2.16\,g \bullet \frac{1\,mol\,Kr}{83.80\,g\,Kr} = 0.0258\,mol\,Kr$$

$$0.0258\,mol\,Kr \bullet \frac{22.4\,L}{1\,mol} = 0.577\,L$$

Gas Stoichiometry

By an extension of Avogadro's Principle, when gases react, coefficients in the balanced equation represent molar amounts AND relative volumes.

Example: $2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)}$

- 2 moles hydrogen gas will react with 1 mole of oxygen gas to produce 2 moles water vapor.
- 2 volumes hydrogen gas will react with 1 volume of oxygen gas to produce 2 volumes water vapor.

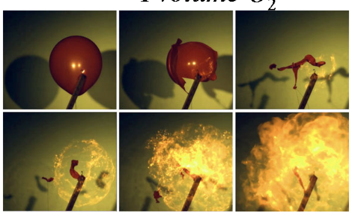
Gas laws can be used to calculate the stoichiometry of reactions where gases are reactants or products.

3. Volume-Volume Example A

How many liters of H_2 gas will react with 5.00 L of O_2 to form water?

$$2\,H_{2(g)} + O_{2(g)} \longrightarrow 2\,H_2O_{(g)}$$

$$5.00\;Liters\,O_2 \bullet \frac{2\;volumes\;H_2}{1\;volume\;O_2} = 10.0\,L\;H_2$$



4. Volume-Volume Example B

How many liters of water will be produced through the combustion of 15.6 L of propane?

$\_\_\; C_3H_{8(g)} + \underline{\;5\;}\; O_{2(g)} \longrightarrow \underline{\;3\;}\; CO_{2(g)} + \underline{\;4\;}\; H_2O_{(g)}$

$$15.6\;L\;\;C_3H_8 \bullet \frac{4\;volumes\;H_2O}{1\;volume\;C_3H_8} = 62.4\,L\;\;H_2$$

5. Volume-Mass Example (Slide 1)

If 7.00 L of N_2 reacts with H_2 at 298 K (P = 3.00 atm), what mass of ammonia is produced?

$$N_{2(g)} + 3\,H_{2(g)} \longrightarrow 2\,NH_{3(g)}$$

Step 1: Determine liters of gaseous ammonia made from 7.00 L of nitrogen gas.

$$7.00\,L\,N_2 \bullet \frac{2\,volumes\,NH_3}{1\,volumes\,N_2} = 14.0\,L\,NH_3$$

5. Volume-Mass Example (Slide 2)

7.00 L of N_2 produces 14.0 L of ammonia at 298 K and pressure of 3.00 atm.

Step 2: Use Ideal Gas Law to find moles.

Data:

V_{NH3} = 14.0 L

P = 3.00 atm

T = 298K

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

$$n = \frac{(3.00\,atm)(14.0\,L)}{\left(0.0821\frac{L\bullet atm}{K\bullet mol}\right)(298\,K)}$$

$$= 1.72\,mol\,NH_3$$

5. Volume-Mass Example (Slide 3)

Finally, find ammonia's molar mass, then make a moles to mass conversion.

Molar mass = 17.3 g/mol.

$$1.72\,mol\,NH_3 \bullet \frac{17.04\,g\,NH_3}{1\,mol\,NH_3} = 29.3\,g\,NH_3$$

The Ideal Gas Law and Density

Remember that density equals mass over volume:

$\rho\,(or\,D) = \frac{m}{V}$	m = mass (g) V = volume (L)

Using molar mass, the equation of gas density is:

$\rho = \frac{MP}{RT}$	M = molar mass (g/mol) P = pressure (atm) R = 0.0821 L atm/K mol T = temperature (K)

Why do we care about density of gases?

Fire Extinguishers! Demo.

6. Density Example

Determine the density of chlorine gas at 22.0 ºC and 1.00 atm.

List known values:

M = 70.90 g/mol
P = 1.00 atm
T = 22.0 ºC = 295 K

$$\rho = \frac{MP}{RT} = \frac{70.90\frac{g}{mol}\bullet 1.00\,atm}{0.0821\frac{L\bullet atm}{K\bullet mol}\bullet 295\,K} = 2.93\frac{g}{L}\,Cl_2$$

Homework:

Read 5.6 in your book.

5.4 Booklet Problems
Due: Next Class