2.A.4 - Gas Laws Part 2: Variable Amount

Lots and Lots of Gas

Chemistry Review! Mole Calculations

Remember what a mole is?

1 mole = Avogadro's number = 6.02 E 23 particles
(in this Unit - gas atoms or molecules)

Mole Calculations

Finding moles of gas from mass:

$n = \frac{m}{M}$	n = moles (mol) m = sample mass (g) M = molar mass (g/mol)

To find M:
1. From gas' formula, make atom inventory.
2. Use Periodic Table and add element masses (in grams) of your inventory.

1. Mole Calculation Example

How many moles of oxygen gas are in 105 grams of oxygen? Oxygen's formula is O_2.

1. Molar mass of oxygen:
   A. Inventory: two atoms of oxygen in formula.
   B. From Periodic Table: Oxygen = 16.0 g/mol.
Molar mass of oxygen: 2 X 16.0 g/mol = 32.0 g/mol.

$$n = \frac{m}{M} = \frac{105\,g}{32.0\frac{g}{mol}} = 3.28\,moles$$

The Ideal Gas Law

So far you've dealt with a fixed amount of gas when making calculations.
When you alter the number of particles of gas (microscopic), or the amount of moles (macroscopic), you have to use the Ideal Gas Law.

Depending on the size of your system, you'll use either the microscopic OR macroscopic form.

Ideal Gas Law - Microscopic

In a small system: gas sample measured in particles:

$PV = Nk_BT$	P = absolute pressure (Pascals (Pa)) V = volume (m^3) N = number of particles k_B= Boltzmann constant (1.38 E -23 J/K) T = temperature (K)

Why measure in particles?
This factors in later, where kinetic energy *per particle* is compared to the internal energy of a system.

Ideal Gas Law - Macroscopic

If your system is large, where your gas sample is measured in moles:

$PV = nRT$	P = absolute pressure (Pascals (Pa)) V = volume (m^3) n = number of moles (mol) R = Universal Gas Constant (8.31 J/(mol•K)) T = temperature (K)

Note: from the microscopic application, when 6.02 E 23 particles (1 mole) is plugged in for N, you get the Universal Gas Constant:

6.02 E 23 particles/mol • 1.38 E -23 J/K = **8.31 J/K•mol**.

2. Macroscopic Example A

What volume will 2.3 moles of an ideal gas at 150 °C occupy at atmospheric pressure?
Remember: atmospheric pressure is 1.0 E 5 Pa.

A Gassy Balloon

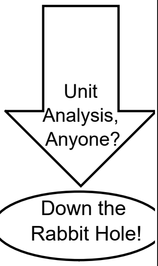
2. Macroscopic Example A Answer

Given:
n = 2.3 moles
P = 1.0 E 5 Pa.
T = 150 °C = CONVERT TO KELVINS!!! = 423 K

$$PV = nRT$$

$$V = \frac{nRT}{P} = \frac{2.3\,mol \bullet 8.31 \frac{J}{K \cdot mol} \bullet 423\,K}{1.0E5\,Pa}$$

$$V = 0.081\,m^3$$



3. Macroscopic Example B

How many grams of helium (molar mass = 4.0 g/mol) does it take to fill a balloon of 15 L volume at STP?
Hint 1: Two-step problem: find moles first.
Hint 2: Convert L to m^3 (1000 L = 1 m^3).

3. Macroscopic Example B Answer
Part 1 - Moles

Given:
V = 15 L = 0.015 m^3
T = 273 K
P = 1.0 E 5 Pa

$$PV = nRT$$

$$n = \frac{PV}{RT} = \frac{1.0E5\,Pa \bullet 0.015\,m^3}{8.31\,J/K \bullet mol \bullet 273\,K} = 0.66\,mol\,He$$

3. Macroscopic Example B Answer
Part 2 - Mass

Given:
molar mass = 4.0 g/mol
moles He = 0.66 mol

$$n = \frac{m}{M}$$

$$m = n \bullet M =$$

$$0.66\,mol \bullet 4.0\,g/mol = 2.64\,grams\,He$$

Non-Ideal Gases

Ideal gases only exist under certain conditions

At high pressure, gases interact with each other

- Intermolecular attractions
- Condensation
- Collisions transfer energy
- Molecular volume effects

Homework

Preview 2.A.5

2.A.4 Booklet Problems
Due: Next Class

Ex. 1 Rabbit Hole - Unit Analysis

How do you get m^3 from the glob of units in the problem?

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

$$= \frac{mol \bullet \frac{J}{K \cdot mol} \bullet K}{Pa} = \frac{J}{Pa} = \frac{N \bullet m}{\frac{N}{m^2}}$$

$$= \frac{N \bullet m}{1} \bullet \frac{m^2}{N} = m^3$$

Back to Reality