Ideal Gas Equation (pg. 479)

Name: SOLUTIONS (5/12/14)

Both pressure and volume depend not only the kinetic energy of the molecules but on their number. We can now combine all four variables that determine the physical characteristics of gases in one equation called the "ideal gas equation". Boyle's law states that pressure is inversely proportional to volume. Charles' law states that volume is directly proportional to temperature. We know that pressure, volume, and temperature all depend on the number of particles present.

$V = k'' \frac{T}{P}$ or $PV = k'' T$, k" is a constant that depends on the number of particles present.

We can write the equation using two constants, n and R, to replace k". **PV = nRT**

n is the number of particles in moles. R is the gas constant whose values can be obtained by substituting standard conditions of P, V, T, and n.

$$R = \frac{PV}{nT} = \frac{(101.325\text{ kPa})(22.4\text{ L})}{(1\text{ mol})(273\text{ K})} = 8.31\ \frac{\text{L kPa}}{\text{mol K}} \quad \text{or} \quad 62.4\ \frac{\text{L mm Hg}}{\text{mol K}} \qquad 0.0821\ \frac{\text{L atm}}{\text{mol K}}$$

Example: What volume will be occupied by 3.25 moles of oxygen gas at .0967 atm and 25°C?

PV = nRT, so $V = \frac{nRT}{P}$

n = 3.25 mol
P = 0.967 atm
T = 25°C + 273 = 298 K
R = 0.0821 L atm/mol K

$$V = \frac{(3.25\text{ mol})(0.0821\text{ L atm})(298\text{ K})}{(0.0967\text{ atm})\quad(\text{mol K})} = 822\text{ L}$$

Since the number of moles (n) of a substance is equal to its mass, (m) divided by its molecular mass (M) the ideal gas equation may be written as:

$PV = \frac{mRT}{M}$. **This equation is useful when the number of mass or molecular weight is required or given.**

Example: In the laboratory, 10.0 g of an unknown gas is found to occupy a volume of 5.60 liters at 20.0°C and 740 mm Hg. What is the molecular mass of this unknown gas?

$PV = \frac{mRT}{M}$, **so:** $M = \frac{mRT}{PV}$

m = 10.0 g **P = 740 mm Hg**
R = 62.4 L mm Hg/mol K **V = 5.60 L**
T = 20°C + 273 = 293 K

$$M = \frac{(10.0\text{ g})(62.4\text{ L mm Hg})(293\text{ K})}{(740\text{ mm Hg})\quad(\text{mol K})(5.60\text{ L})} = 44\text{ g/ mole}$$

The gas laws apply to ideal gas behavior (that is a gas without volume nor attractive forces between its molecules). Under most conditions real gases do behave ideally. However, for many gases at conditions of high pressure and low temperature ideal behavior is not possible do to van der Waals forces.

As pressure is increased, the gas molecules are forced closer together. If the temperature is low enough the molecules will be moving slow and will be affected by van der Waals forces. The molecules will come together by van der Waals attraction and the gas becomes a liquid. If the temperature is high the pressure must be increased further in order to liquify a gas. Above a certain temperature, no amount of pressure will liquify a gas. This temperature is called the critical temperature and is unique for each gas. Gases with low critical temperatures are difficult to liquify and thus approximate ideal gas behavior.

Problems:

1. How many moles of gas will occupy a 486 mL flask at 10.0°C and 66.7 kPa pressure? (ans. 0.0138 mol)

equation: $PV = nRT$

rearranged equation: $n = \frac{PV}{RT}$

equation w/ variables substituted: $n = \frac{(66.7\ \cancel{kPa})(0.486\ \cancel{L})}{(8.31\ \frac{\cancel{kPa \cdot L}}{mol \cdot \cancel{K}})(283\ \cancel{K})}$

$P = 66.7\ kPa$
$V = 486\ mL = 0.486\ L$
$T = 273 + 10 = 283\ K$
$R = 8.31\ \frac{kPa\ L}{mol\ K}$

$\boxed{n = 0.0138\ mol}$

2. What is the molecular mass of a gas if 6.71 g of it occupy 2.12 L at 100.4 kPa and 29°C? (ans. 79.1 g/mol)

"molar mass" (g/mol)

equation:
1) $PV = nRT$
2) $M = \frac{m}{n}$

rearranged equation:
1) $n = PV/RT$
2) $M = \frac{m}{n}$

equation w/ variables substituted:
1) $n = \frac{(100.4\ \cancel{kPa})(2.12\ \cancel{L})}{(8.31\ \frac{\cancel{kPa\ L}}{mol\ \cancel{K}})(29 + 273\ \cancel{K})}$
$= 0.0848\ mol$

2) Molar mass $= \frac{6.71\ g}{0.0848\ mol} = \boxed{79.1\ g/mol}$

3. A sample of gas occupies 14.3 L at 19.0 C and 1.2 atm. How many moles of gas are present? (ans. .72 mol)

equation: $PV = nRT$

rearranged equation: $n = \frac{PV}{RT}$

equation w/ variables substituted: $n = \frac{(1.2\ \cancel{atm})(14.3\ \cancel{L})}{(0.082\ \frac{\cancel{atm\ L}}{mol\ \cancel{K}})(19 + 273\ \cancel{K})}$

$\boxed{n = 0.72\ mol}$

4. How many moles of Hydrogen gas are present in a 50.0 L cylinder if the pressure is 10atm and the temp. is 27.0 C? (R=.082 liter-atm/mol-K) (ans. 20.3 mol)

equation: $PV = nRT$

rearranged equation: $n = \frac{PV}{RT}$

equation w/ variables substituted: $n = \frac{(10\ \cancel{atm})(50.0\ \cancel{L})}{(0.082\ \frac{\cancel{atm\ L}}{mol\ \cancel{K}})(300\ \cancel{K})}$

$\boxed{n = 20.3\ mol\ H_2}$

5. A cylinder contains oxygen at a pressure of 10.0 atm and a temperature of 300K. The volume of the cylinder is 10.0 L. What is the mass of the oxygen? (you will have to obtain answer and then go from M to g) (ans. 130 g)

equation:
1) $PV = nRT$
2) Molar Mass $= \frac{m}{n}$

rearranged equation:
1) $n = \frac{PV}{RT}$
2) $m = M \cdot n$

equation w/ variables substituted:
1) $n = \frac{(10.0\ \cancel{atm})(10.0\ \cancel{L})}{(0.082\ \frac{\cancel{atm\ L}}{mol\ \cancel{K}})(300\ \cancel{K})}$
$= 4.07\ mol\ O_2$

2) $m = 4.07\ \cancel{mol\ O_2} \times \frac{32\ g\ O_2}{1\ \cancel{mol\ O_2}} = \boxed{130\ g\ O_2}$