REAL GASES

- The empirical gas laws (including the ideal gas equation) do not always apply.

    - The gas laws don't apply in situations where the assumptions made by kinetic theory are not valid.

        - When would it be FALSE that the space between gas molecules is much larger than the molecules themselves?

            - at high pressure, molecules would be much closer together!

        - When would it be FALSE that attractive and repulsive forces would be negligible?

            - at high pressure, attractions and repulsions should be stronger!

            - at low temperature, attractions and repulsions have a more significant affect on the paths of molecules

-The gas laws are highly inaccurate near the point where a gas changes to liquid!

- In general, the lower the pressure and the higher the temperature, the more IDEAL a gas behaves.

143 van der Waals equation

- an attempt to modify PV = nRT to account for several facts.
  - gas molecules actually have SIZE (they take up space)
  - attractive and repulsive forces

$$PV = nRT$$ Ideal gas equation

$$\left(P + \frac{n^2 a}{V^2}\right)(V - nb) = nRT$$ van der Waals equation

- $\frac{n^2 a}{V^2}$: attempts to account for attractive / repulsive forces
- nb: attempts to account for molecular size

* "a" and "b" are experimentally determined parameters that are different for each gas. p211

He: a = 0.0346, b = 0.0238 tiny, no special attractive forces

H_2O: a = 5.537, b = 0.03049 small, but strong attractions between molecurles

CH_3CH_2OH: a = 12.56 b = 0.08710 larger, and strong attractions between molecules

2500 L of chlorine gas at 25.0 C and 1.00 atm are used to make hydrochloric acid. How many kilograms of hydrochloric acid could be produced if all the chlorine reacts?

$$H_2 + Cl_2 \rightarrow 2\,HCl$$

1 - Convert 2500L of chlorine gas to moles. Use ideal gas equation.
2 - Convert moles chlorine gas to moles hydrochloric acid using chemical equation
3 - Convert moles hydrochloric acid to mass using formula weight

$PV = nRT$

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

$P = 1.00\ atm$ $\quad T = 25.0°C = 298.2\ K$ $\quad n = ?\ mol$

$R = 0.08206\ \frac{L\cdot atm}{mol\cdot K}$ $\quad V = 2500\ L$

(1) $$n_{Cl_2} = \frac{(1.00\ atm)(2500\ L)}{\left(0.08206\ \frac{L\cdot atm}{mol\cdot K}\right)(298.2\ K)} = 102.164983\ mol\ Cl_2$$

$mol\ Cl_2 = 2\ mol\ HCl$ | $36.458\ g\ HCl = mol\ HCl$ | $kg = 10^3\ g$

$$102.164983\ mol\ Cl_2 \times \frac{2\ mol\ HCl}{mol\ Cl_2} \times \frac{36.458\ g\ HCl}{mol\ HCl} \times \frac{kg}{10^3\ g} = \boxed{7.45\ kg\ HCl}$$

(2) (3)

Calculate the mass of 22650 L* of oxygen gas at 25.0 C and 1.18 atm pressure.

← O_2

*Volume of a 10'x10'x8' room

1 - Convert volume of oxygen gas to moles using ideal gas equation
2 - Convert moles oxygen gas to mass using formula weight.

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

$P = 1.18 \text{ atm}$ $\quad n = ? \text{ mol}$

$V = 22650 \text{ L}$ $\quad T = 25.0°C = 298.2 \text{ K}$

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

(1) $$n_{O_2} = \frac{(1.18 \text{ atm})(22650 \text{ L})}{\left(0.08206 \frac{L \cdot atm}{mol \cdot K}\right)(298.2 \text{ K})} = 1092.222357 \text{ mol } O_2$$

(2) $32.00 \text{ g } O_2 = \text{mol } O_2$

$$1092.222357 \text{ mol } O_2 \times \frac{32.00 \text{ g } O_2}{\text{mol } O_2} = 35\bar{0}00 \text{ g } O_2$$

35.0 kg O_2

OR

77 lb