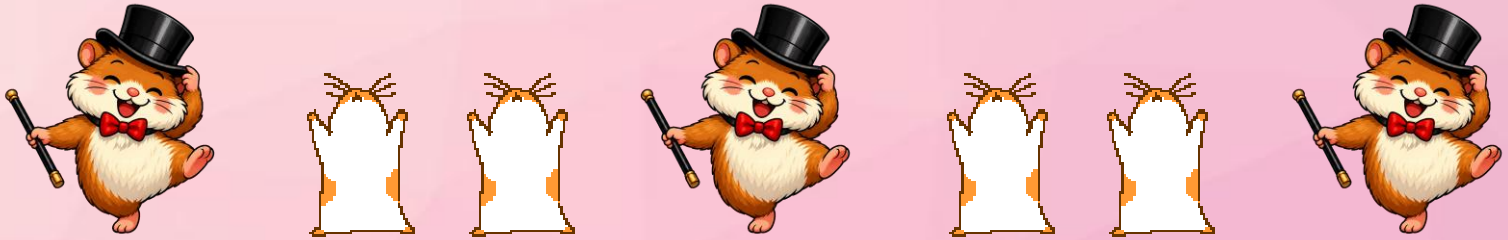

Counting Particles In Substances

Chemistry

Karl Steffin, 2001

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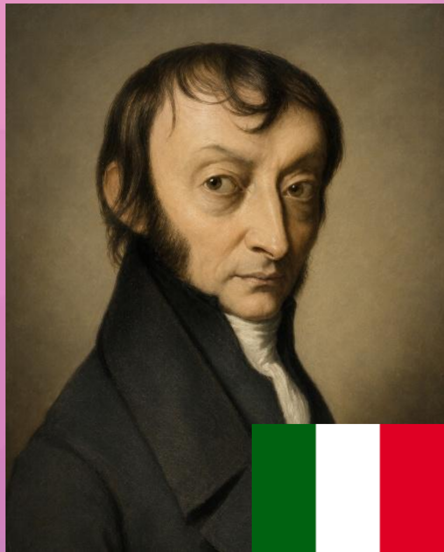
By end of this unit I can…

- A.1 Explain the relationship between the mass of a substance, the number of particles of that substance, and the number of moles of that substance.
- A.2 Use the mole concept to calculate the mass, number of particles, or number of moles of a given substance.
- B.1 Explain the relationships between macroscopic properties of gas samples.
- B.2 Perform calculations using the ideal gas law.
- B.3 Create and/or evaluate models based on the ideal gas law.

Amadeo Avogadro Di Quaregna

- **Avogadro stated that 1 mole (n) is equal to 602,000,000,000,000,000,000,000 particles.**
  - **Avagadro's Number:** $6.02x10^{23}\frac{1}{mol}$
  - **n·N_A= N**



What does this mean?

- Moles (n) are a universal 'middleman' to convert to/from other units.
  - Ex: Calculate how many grams are in one mole of Potassium Hydroxide (KOH).
- The number of grams needed is the sum of the periodic weights of K, O, and H.
  - Finding periodic weights: In Class always round to the tenth.

39.1-g + 16.0-g + 1.0-g = 56.1-g

1-mole KOH = $\boxed{\text{56.1-grams KOH}}$

Trial and Error

- **How many grams are in 1 mole of H_2SO_4?**

  2 H (1.0-g each) = 2.0-g

  1 S (32.1-g) = 32.1-g

  4 O (16.0-g each) = 64.0-g

  2.0-g + 32.1-g + 64.0-g = 98.1-g

  1-mole H_2SO_4 = $\boxed{98.1\text{-g } H_2SO_4}$

Trial and Error Cont.

- How many grams are in 3 moles of H_2O?

Given 2 H (1.0-g each) and 1 O (16.0-g).

$$2(1.0\text{-g}) + 16.0\text{-g} = 18.0\text{-g}$$

One mole is 18.0-g, but three are needed…

$$18.0\text{-g} \times 3 = 54.0\text{-g}$$

$$3\text{-moles } H_2O = \boxed{54.0\text{-g } H_2O.}$$

Error and Trial

- How many moles are in 702.0-g of NaCl?

One mole of NaCl is (23.0-g+35.5-g=58.5-g)

Truth for NaCl: 1-mole = 58.5-g

To get rid of the mass divide.

$$702.0\text{-g NaCl} \times \frac{1\text{-mole}}{58.5\text{-g}} = \boxed{12.0\text{-moles NaCl}}$$

702.0-g NaCl = 12.0-mole NaCl.

Gasses: Micro to Macro

- **Micro View:** $\frac{PV}{NT} = k_b$ $\quad k_b \approx 1.38x10^{-23}\frac{J}{K}$ $\quad n \cdot N_A = N$
- **To look at this in a macroscopic way:**

$$\frac{PV}{nN_AT} = k_b \qquad \frac{PV}{nT} = k_bN_A$$

- $k_bN_A \approx (1.38x10^{-23}\frac{J}{K})(6.02x10^{23}\frac{1}{mol}) = 8.314 - \frac{J}{mol \cdot K}$
- $8.314 - \frac{J}{mol \cdot K} = Ideal\ Gas\ Constant: R$

Gasses: Ideal Gas Law

- **Macroscopic view gives us the Ideal Gas Law: PV=nRT**
  - **Ideal Gas Constant:** $\mathrm{R} = 8.314 - \frac{J}{mol \cdot K}$

Energy→ $1 - J = 1 - N \cdot m$

Pressure→ $1 - Pa = 1 - \frac{N}{m^2}$ **or** $1 - Pa \cdot m^2 = 1 - \mathrm{N}$

Substitute→ $1 - J = 1 - Pa \cdot m^2 \cdot m = 1 - Pa \cdot m^3$

Pressure→ $1 - atm = 101325 - \mathrm{Pa}$

$$1 - Pa = \frac{1}{101325} - \mathrm{atm}$$

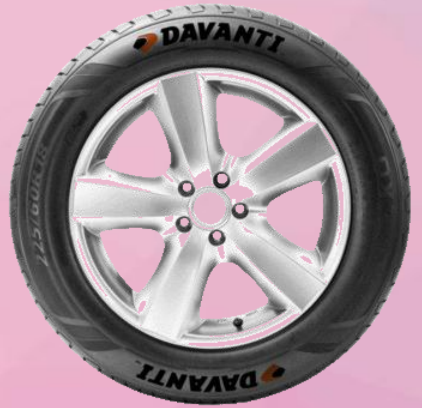
Combine II→ $1 - J = \frac{1}{101325} - atm \cdot 1000 - L$

Volume→ $1 - m^3 = 1000 - \mathrm{L}$

Combine III→ $R = \frac{8.314 \cdot 1000}{101325} - \frac{atm \cdot L}{mol \cdot K}$

$$R = .0821 - \frac{atm \cdot L}{mol \cdot K}$$

Gas Law Example: Ideal



- In a tire what volume does 6-moles of N_2 occupy at 1.5-atm and 16-°C?

$$PV = nRT$$

$$1.5 - \cancel{atm} \cdot V = 6 - \cancel{mol} \cdot .0821 \frac{\cancel{atm} \cdot L}{\cancel{mol} \cdot \cancel{K}} \cdot 289.15 - \cancel{K}$$

$$1.5\,V = 142.43529 - L$$

$$V = 94.95686 - L$$

$P = 1.5 - atm$

$V =$

$n = 6 - mol$

$R = .0821 - \frac{atm \cdot L}{mol \cdot K}$

$T = 16 - ℃$

$T = 289.15 - K$

$\boxed{95.0\text{-}L}$

Gasses: Molar Standard

- To create a repeatable condition a standard Pressure and Temperature have been established for gasses.
  - Pressure: 1-atm Temperature: 273.15-K
- To set up a molar ratio assume n=1-mol

$$PV = nRT$$

$$1-atm \cdot V = 1-mol\left(.0821-\frac{atm \cdot L}{mol \cdot K}\right)273.15-K$$

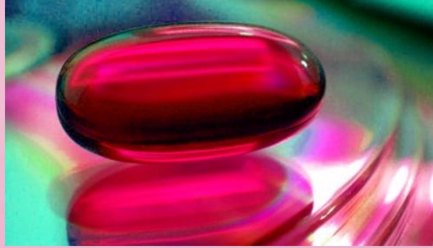
$$V = (.0821)273.15-L$$

$$V = 22.4-L$$

At STP: 1-mol=22.4-L

$$P = 1-atm$$
$$V =$$
$$n = 1-mol$$
$$R = .0821-\frac{atm \cdot L}{mol \cdot K}$$
$$T = 273.15-K$$

The 'Mole' Matrix



Going to Moles		Leaving Moles
Volume → $\frac{1\text{-mol}}{22.4\text{-L}}$	→ Mole →	$\frac{22.4\text{-L}}{1\text{-mol}}$ → Volume
Mass → $\frac{1\text{-mol}}{\text{P. Table}}$	→ Mole →	$\frac{\text{P. Table}}{1\text{-mol}}$ → Mass
Particles → $\frac{1\text{-mol}}{N_{Ava}}$	→ Mole →	$\frac{N_{Ava}}{1\text{-mol}}$ → Particles

Using the Matrix 1 Step

How many Liters are in 15.0-moles of Nitrogen Dioxide at STP?

$$15.0\text{-}\cancel{\text{mol}}\ NO_2 \times \frac{22.4\text{-L}}{1\text{-}\cancel{\text{mol}}} = \boxed{336.0\text{-L}\ NO_2}$$

Using the Matrix 2 Steps

How many molecules are in 250.0-g of Table Sugar ($C_{12}H_{22}O_{11}$)?

$$250.0\text{-}\cancel{g}\ \text{Sugar} \times \frac{1\text{-}\cancel{mol}}{342\text{-}\cancel{g}} \times \frac{N_a\text{-Mol}}{1\text{-}\cancel{mol}}$$

$$\boxed{= 4.4 \times 10^{23}\text{-Mol Sugar}}$$

mol is mole while Mol is Molecule

1 Mol sugar has 45 atoms in it (another truth)

Empirical Formulas

EF: A formula that gives the simple whole number ratio of each element.

1. **Find the Molar Mass of each Element.**
2. **Divide each ans. by the smallest ans.**
3. **If any answer in step 2 is a fraction: multiply each by the inverse fraction.**

Ex) Determine the Empirical Formula for a compound containing 1.203-g Ca and 2.128-g Cl.

Empirical Formulas EX I

1. Find the Molar Mass of each Element.

- Ca: 1.203-g x 1-mol/40.1-g= .030-mol
- Cl: 2.128-g x 1-mol/35.5-g= .060-mol

2. Divide each by the smallest number.

- Ca: .030/.030=1
- Cl: .060/.030=2
- For every 1 Ca there are 2 Cl's.
  - Step three was not needed in this problem.

$\boxed{CaCl_2}$

Empirical Formulas EX II

What is the EF of a compound containing 77.72% Iron, 22.28% Carbon?

- **If given %, just change the unit to grams.**

1. **77.72-g Fe x 1-mol/55.8-g= 1.3928 Fe**

   22.28-g C x 1-mol/12.0-g= 1.8566 C
2. **Fe: 1.3928/1.3928 = 1**

   C: 1.8566/1.3928 = 1.333 (1 ⅓)
3. **Multiply by the fractions (⅓) inverse.**

   Fe: 1 x 3 = 3 , C: 1.333 x 3 = 4

$\boxed{Fe_3C_4}$

Determining Molecular Formulas

- Empirical models only give ratios
- There is a huge difference between the ion OH^- and the molecule H_2O_2.
  - Both have a 1:1 ratio though
- Determining the MF that considers the actual number of each atom in a compound may be more beneficial.

Molecular Form Example I

Find the MF for a compound that contains 4.90-g N and 11.2-g O. The compounds Molar Mass is 92.0 g/mol.

- **First find the Empirical Formula:**
  - **N: 4.90-g x 1 mol/14.0-g = .35 .35/.35 = 1**
  - **O: 11.2-g x 1 mol/16.0-g = .70 .70/.35 = 2**

- **EF: NO_2 so the MF is any ratio of $N_{(x)}O_{(2x)}$**

Molecular Form Example I cont.

- The smallest mass of $N_{(x)}O_{(2x)}$ is x=1 or NO_2.
- The Mass of NO_2 is 46.0-g.
- x = MF Mass / EF Mass

$$\frac{92.0\text{-g}}{46.0\text{-g}} = 2 \text{ (will always be an integer)}$$

- Plug back into x. $\boxed{N_2O_4}$

Molecular Form Example II

β-carotene (536-g/mol) has an EF of C_5H_7. Find the Molecular Formula.

- We have the Empirical Formula so:
  - MF = $C_{5x}H_{7x}$
  - EF (x=1): C_5H_7 = 67-g
  - X=MF Mass/EF Mass

$$\frac{536\text{-g}}{67\text{-g}} = 8$$

$$\boxed{C_{40}H_{56}}$$