Chemical Composition

Atomic Weight

Atoms are so tiny.

$\Downarrow$

We use a new unit of mass:

Atomic mass unit (amu) = 1.6605×10^{-24} g

Atomic Weight (Mass)

Atomic weight (mass): of an element is average of the masses (in amu) of its isotopes found on the Earth.

Cl

$^{35}_{17}Cl \longrightarrow$ 34.97 amu

$^{37}_{17}Cl \longrightarrow$ 36.97 amu

$$(75.77/100 \times 34.97 \text{ amu}) + (24.23/100 \times 36.97 \text{ amu}) = 35.45 \text{ amu}$$

$$\begin{array}{lc} \text{Atomic number} \longrightarrow & 17 \\ & \text{Cl} \\ \text{Atomic weight} \longrightarrow & 35.45 \end{array}$$

Counting atoms by weighing

Say we need 1000 of each.

Who wants to count?

Counting atoms by weighing

- Often it is useful to *count* the number of items by *weighing* them instead.

How many Jelly beans in 60.2 g?

$$60.2\ \text{g} \times \frac{1\ \text{JB}}{3.01\ \text{g}} = 20\ \text{JB}$$

How many Mint tic tac in 44.1 g?

$$44.1\ \text{g} \times \frac{1\ \text{M}t}{1.26\ \text{g}} = 35\ \text{Mt}$$

Jelly Bean (g)	Mint (g)
3.01	1.03
2.89	1.50
3.20	1.30
3.10	1.30
2.95	1.20
2.97	1.10
3.02	1.25
3.05	1.30
2.99	1.43
2.91	1.15
3.01	1.26

Counting atoms by weighing

5.03×10^{30} amu Cl = ? Atoms of Cl

1 Chlorine atom = 35.45 amu

$$5.03 \times 10^{30} \cancel{\text{ amu Cl}} \times \frac{1 \text{ Cl atom}}{35.45 \cancel{\text{ amu Cl}}} = 1.42 \times 10^{29} \text{ atoms Cl}$$

Atomic mass (weight) $\longleftrightarrow$ Number of atoms

Formula and Molecule

Ionic & covalent compounds $\rightarrow$ Formula formula of NaCl

Covalent compounds $\rightarrow$ Molecule molecule of H_2O

Formula Weight of NaCl:

23 amu Na + 35.5 amu Cl = 58.5 amu NaCl

Molecular Weight of H_2O:

$2 \times (1 \text{ amu H}) + 16 \text{ amu O} = 18 \text{ amu } H_2O$

Mole

Mole: formula weight of a substance (in gram).

12g of C = 1 mol C

23g of Na = 1 mol Na

58.5 g of NaCl = 1 mol NaCl

18 g of H_2O = 1 mol of H_2O

1 dozen = 12 of something.

1 mole = 6.022×10^{23} of something.

Avogadro's number (6.02×10^{23}): number of formula units in one mole.

1 mole of apples = 6.02×10^{23} apples

1 mole of A atoms = 6.02×10^{23} atoms of A

1 mole of A molecules = 6.02×10^{23} molecules of A

1 mole of A ions = 6.02×10^{23} ions of A

Molar Mass

Molar mass (g/mol): mass of 1 mole of substance (in gram).
(Formula weight)

molar mass of Na = 23 g/mol
molar mass of H_2O = 18 g/mol

Calculation of moles & number of molecules

$$104 \text{ g of NaCl} \longrightarrow \begin{array}{l} \text{? moles} \\ \text{? molecules} \end{array}$$

Formula weight = 1×(23 amu) for Na + 1×(35.5 amu) for Cl = 58.5 amu NaCl

1 mole NaCl = 58.5 g NaCl

$$104 \cancel{\text{ g NaCl}} \times \frac{1 \text{ mole NaCl}}{58.5 \cancel{\text{ g NaCl}}} = 1.78 \text{ moles NaCl}$$

1 mole NaCl = 6.02×10^{23} molecules NaCl

$$1.78 \cancel{\text{ moles NaCl}} \times \frac{6.02 \times 10^{23} \text{ molecules NaCl}}{1 \cancel{\text{ mole NaCl}}} = 1.07 \times 10^{24} \text{ molecules NaCl}$$

Stoichiometry

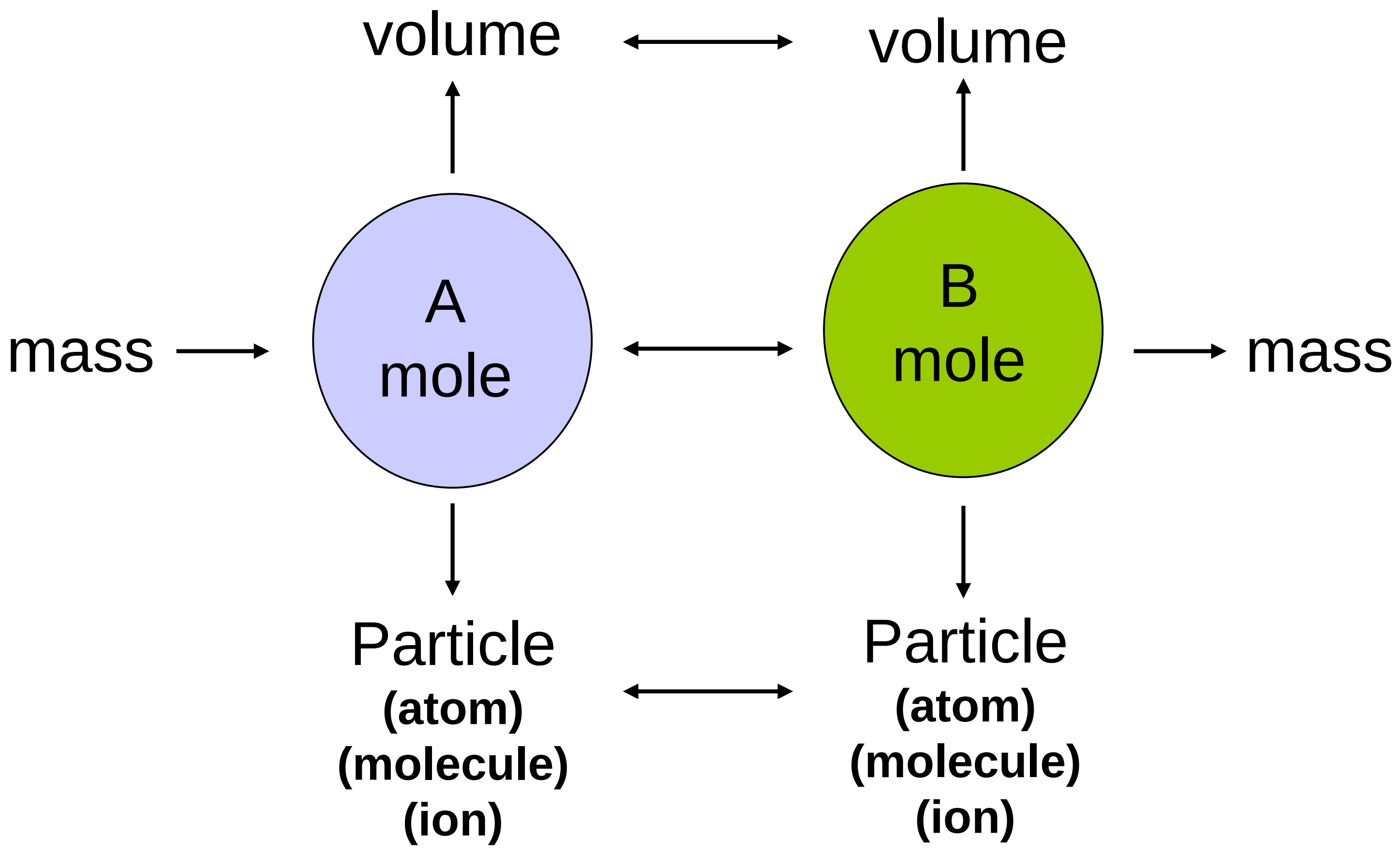
Relationships between amounts of substances in a chemical reaction.

Look at the Coefficients!

$$2H_2O_{(l)} \rightarrow 2H_{2(g)} + O_{2(g)}$$

2	2	1
2 moles	2 moles	1 mole
2 liters	2 liters	1 liter
2 particles	2 particles	1 particle
~~2 grams~~	~~2 grams~~	~~1 gram~~

Stoichiometry



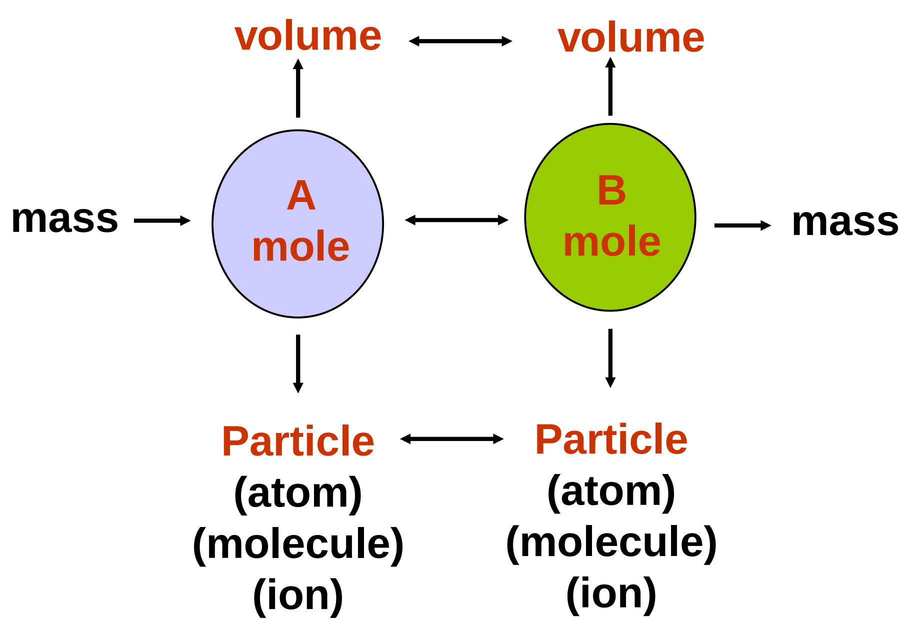
1 Step

Mole A ⇔ Mole B

Volume A ⇔ Volume B

# of Particles A ⇔ # of Particles B

Use coefficient in the balanced equation



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

$$\overset{A}{23 \text{ mole } CH_4} = ? \overset{B}{\text{ moles } H_2O}$$

$$\cancel{23 \text{ mole } CH_4} \left(\frac{2 \text{ moles } H_2O}{\cancel{1 \text{ mole } CH_4}} \right) = 46 \text{ moles } H_2O$$

$$\overset{A}{10 \text{ L } O_2} = ? \overset{B}{\text{ L } CO_2}$$

$$10 \cancel{\text{ L } O_2} \left(\frac{1 \text{ L } CO_2}{2 \cancel{\text{ L } O_2}} \right) = 5 \text{ L } CO_2$$

$$\overset{A}{2\times10^{26} \text{ molecules } H_2O} = ? \overset{B}{\text{ molecules } O_2}$$

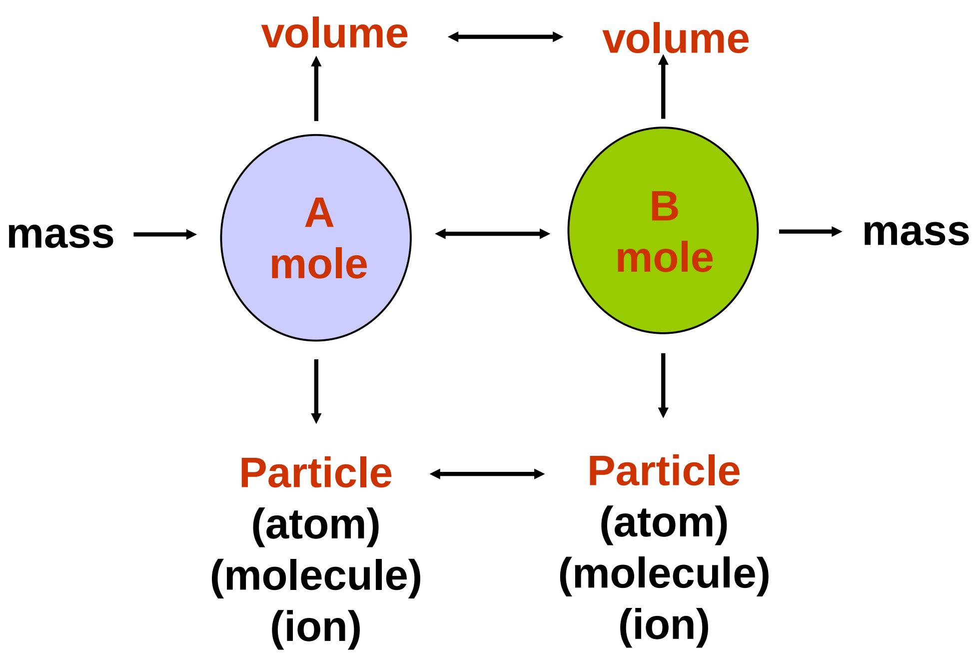
$$2\times10^{26} \cancel{\text{ molecules } H_2O} \left(\frac{2 \times (6.02\times10^{23} \text{ molecules } O_2)}{2 \times (6.02\times10^{23} \cancel{\text{ molecules } H_2O})} \right) = 2\times10^{26} \text{ molecules } O_2$$

2 Steps

Mole A ⇔ Volume B

Mass A ⇔ Mole B or Volume A

# of Particles A ⇔ Mole B or Volume A



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

$$\overset{A}{32 \text{ g } CH_4} = ? \text{ moles } \overset{B}{CO_2}$$

$$32 \text{ g } CH_4 \left(\frac{1 \text{ mole } CH_4}{16 \text{ g } CH_4}\right)\left(\frac{1 \text{ mole } CO_2}{1 \text{ mole } CH_4}\right) = 2.0 \text{ mole } CO_2$$

$$\overset{A}{40. \text{ g } CH_4} = ? \text{ L } \overset{A}{CH_4}$$

$$40. \text{ g } CH_4 \left(\frac{1 \text{ mole } CH_4}{16 \text{ g } CH_4}\right)\left(\frac{22.4 \text{ L } CH_4}{1 \text{ mole } CH_4}\right) = 56 \text{ L } CH_4$$

STP: **1 mole of substance (gas) = 22.4 L = 22400 cc (cm^3 or mL)**

$$\overset{A}{5 \text{ moles } CO_2} = ? \text{ molecules } \overset{B}{O_2}$$

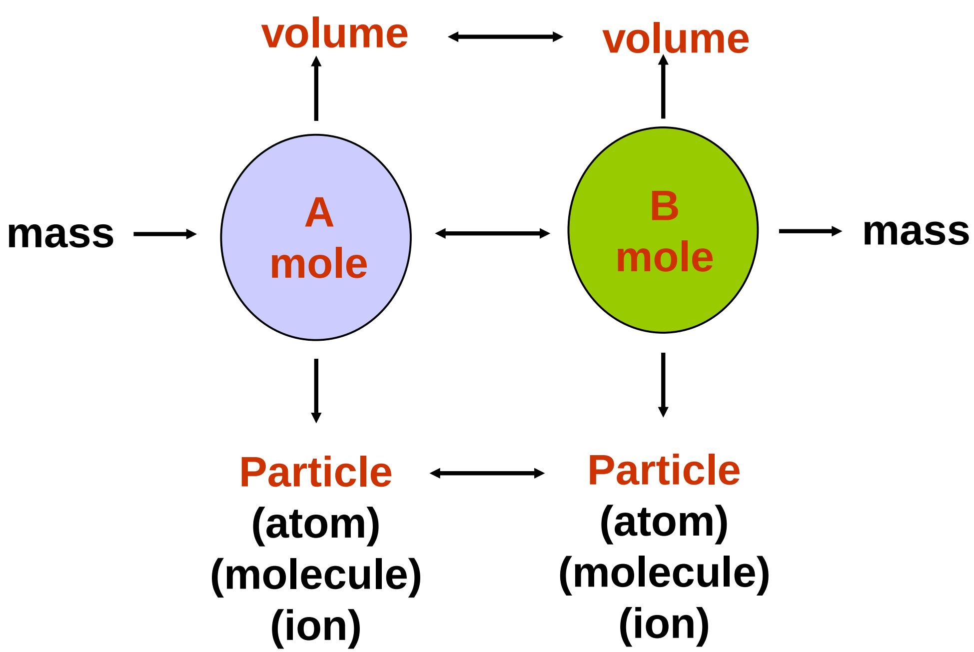
$$5 \text{ moles } CO_2 \left(\frac{2 \text{ mole } O_2}{1 \text{ mole } CO_2}\right)\left(\frac{6.02\times10^{23} \text{ molecules } O_2}{1 \text{ mole } O_2}\right) = 6\times10^{24} \text{ molecules } O_2$$

3 Steps

Mass A ⇔ Mass B

Mass A ⇔ Volume B or # of Particles B

# of Particles A ⇔ Volume B



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

A B

46.0 g CH_4 = ? g H_2O

$$46.0\ \text{g}\ CH_4 \left(\frac{1\ \text{mole}\ CH_4}{16\ \text{g}\ CH_4}\right)\left(\frac{2\ \text{mole}\ H_2O}{1\ \text{mole}\ CH_4}\right)\left(\frac{18\ \text{g}\ H_2O}{1\ \text{mole}\ H_2O}\right) = 104\ \text{g}\ H_2O$$

Mass Percent

$$\text{Percent} = \frac{\text{Part}}{\text{Whole}} \times 100$$

$$\text{Mass Percent of element} = \frac{\text{Mass of element in 1 mole compound}}{\text{Mass of 1 mole compound}} \times 100$$

Mass Percent

$C_6H_{12}O_6$ Mass percent of C, O, H = ? %

$$\text{Mass of C} = 6\ \cancel{\text{mol}} \times 12.01\ \text{g}/\cancel{\text{mol}} = 72.06\ \text{g}$$

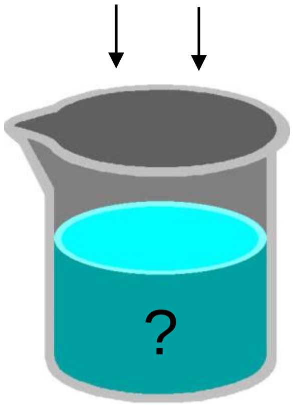
$$\text{Mass of O} = 6\ \cancel{\text{mol}} \times 16.00\ \text{g}/\cancel{\text{mol}} = 96.00\ \text{g}$$

$$\text{Mass of H} = 12\ \cancel{\text{mol}} \times 1.008\ \text{g}/\cancel{\text{mol}} = 12.09\ \text{g}$$

$$\text{Mass of 1 mole } C_6H_{12}O_6 = 180.16\ \text{g}$$

$$\left.\begin{aligned}
\text{Mass percent of C} &= \frac{72.06\ \text{g}}{180.16\ \text{g}} \times 100 = 40.00\% \\
\text{Mass percent of O} &= \frac{96.00\ \text{g}}{180.16\ \text{g}} \times 100 = 53.00\% \\
\text{Mass percent of H} &= \frac{12.09\ \text{g}}{180.16\ \text{g}} \times 100 = 6.700\%
\end{aligned}\right\} + = 100\%$$

Finding formulas of compounds



C = 0.0806 g
O = 0.1074 g
H = 0.01353 g

$\Rightarrow$ Formula of compound?

$$0.0806\ \cancel{g\ C} \times \frac{1 \text{ mole C atoms}}{12.01\ \cancel{g\ C}} = 0.00671 \text{ mol C atoms}$$

$$0.1074\ \cancel{g\ O} \times \frac{1 \text{ mole O atoms}}{16.00\ \cancel{g\ O}} = 0.006713 \text{ mol O atoms}$$

$$0.01353\ \cancel{g\ H} \times \frac{1 \text{ mole H atoms}}{1.008\ \cancel{g\ H}} = 0.01342 \text{ mol H atoms}$$

Finding formula of compounds

6.02×10^{23} Formula unit in 1 mol

$$0.00671\ \cancel{\text{mol C atoms}} \times \frac{6.02\times10^{23}\ \text{C atoms}}{\cancel{1\ \text{mol C atoms}}} = 4.04\times10^{21}\ \text{C atoms}$$

$$0.006713\ \cancel{\text{mol O atoms}} \times \frac{6.02\times10^{23}\ \text{O atoms}}{\cancel{1\ \text{mol O atoms}}} = 4.040\times10^{21}\ \text{O atoms}$$

Equal

$$0.01342\ \cancel{\text{mol H atoms}} \times \frac{6.02\times10^{23}\ \text{H atoms}}{\cancel{1\ \text{mol H atoms}}} = 8.080\times10^{21}\ \text{H atoms} \longrightarrow \text{twice}$$

CH_2O or $C_2H_4O_2$ or $C_3H_6O_3$

1:2:1

Empirical formula

1:2:1

Relative numbers of atoms

Smallest whole-number ratio

CH_2O

Simplest formula (Empirical formula)

Calculation of Empirical formula

Step 1: Find the mass of each element (in grams).

Step 2: Determine the numbers of moles of each type of atom present (using atomic mass).

Step 3: Divide the number of moles of each element by the smallest number of moles to convert the smallest number to 1.

If all of numbers so obtained are integers (whole numbers), these numbers are the subscripts in the empirical formula. If no, we go to step 4:

Step 4: Multiply the numbers that you obtained in step 3 by the smallest integer that will convert all of them to whole numbers (always between 1 and 6).

Calculation of Empirical formula

98.55% Ba

$\longrightarrow$ Empirical formula ?

1.447% H

Consider 100.00g of compound:

$$98.55 \text{ g } \cancel{\text{Ba}} \times \frac{1 \text{ mole Ba atoms}}{137.3 \text{ g } \cancel{\text{Ba}}} = 0.7178 \text{ mol Ba atoms}$$

$$1.447 \text{ g } \cancel{\text{H}} \times \frac{1 \text{ mole H atoms}}{1.008 \text{ g } \cancel{\text{H}}} = 1.436 \text{ mol H atoms}$$

$$\frac{0.7178 \text{ mol Ba}}{0.7178} = 1 \text{ mol Ba}$$

Integers

$$\boxed{BaH_2}$$

$$\frac{1.436 \text{ mole H}}{0.7178} = 2 \text{ mol H}$$

Empirical formula

Calculation of Empirical formula

4.151 g Al

$\longrightarrow$ Empirical formula ?

3.692 g O

$$4.151\ \cancel{\text{g Al}} \times \frac{1\ \text{mole Al atoms}}{26.98\ \cancel{\text{g Al}}} = 0.1539\ \text{mol Al atoms}$$

$$3.692\ \cancel{\text{g O}} \times \frac{1\ \text{mole O atoms}}{16.00\ \cancel{\text{g O}}} = 0.2308\ \text{mol O atoms}$$

$$\frac{0.1539\ \text{mol Al}}{0.1539} = 1\ \text{mol Al}$$

$$\frac{0.2308\ \text{mole O}}{0.1539} = 1.5\ \text{mol O}$$ is not integer

Calculation of Empirical formula

$$1.00 \text{ Al} \times 2 = 2 \text{ Al atoms}$$

$$1.50 \text{ O} \times 2 = 3 \text{ O atoms}$$

$$\boxed{Al_2O_3}$$

Empirical formula

Empirical formula & Molecular formula

Molecular formula = (empirical formula)$_n$

$$C_6H_{12}O_6 = (CH_2O)_6$$

Molecular formula = n $\times$ empirical formula

Molar mass = n $\times$ empirical formula mass

$$\boxed{n = \frac{\text{Molar mass}}{\text{Empirical formula mass}}}$$

Empirical formula & Molecular formula

71.65% Cl

24.27% C

4.07% H

Molar mass = 98.96 g/mol $\longrightarrow$ Empirical formula ? Molecular formula ?

Consider 100.00g of compound:

$$71.65 \text{ g } \cancel{\text{Cl}} \times \frac{1 \text{ mole Cl atoms}}{35.45 \text{ g } \cancel{\text{Cl}}} = 2.021 \text{ mol Cl atoms}$$

$$24.27 \text{ g } \cancel{\text{C}} \times \frac{1 \text{ mole C atoms}}{12.01 \text{ g } \cancel{\text{C}}} = 2.021 \text{ mol C atoms}$$

$$4.07 \text{ g } \cancel{\text{H}} \times \frac{1 \text{ mole H atoms}}{1.008 \text{ g } \cancel{\text{H}}} = 4.04 \text{ mol H atoms}$$

Empirical formula & Molecular formula

$$\frac{2.021 \text{ mol Cl}}{2.021} = 1 \text{ mol Cl}$$

$$\frac{2.021 \text{ mole C}}{2.021} = 1 \text{ mol C}$$

$$\frac{4.04 \text{ mol H}}{2.021} = 2 \text{ mol H}$$

$\boxed{ClCH_2}$ Empirical formula

$$n = \frac{\text{Molar mass}}{\text{Empirical formula mass}}$$

$n = \frac{98.96g}{48.468g} = 2$ $(ClCH_2)_2$ $\boxed{Cl_2C_2H_4}$ Molecular formula

At-Home Practice

Practice problem:

A compound has a molar mass between 165 – 170 g. The percentage composition by mass values are carbon, 42.87%; hydrogen, 3.598%; oxygen, 28.55%; nitrogen, 25.00%. Determine the *empirical formula* and the *molecular formula* of the compound.

- What is the empirical formula?
- What is the molecular formula?