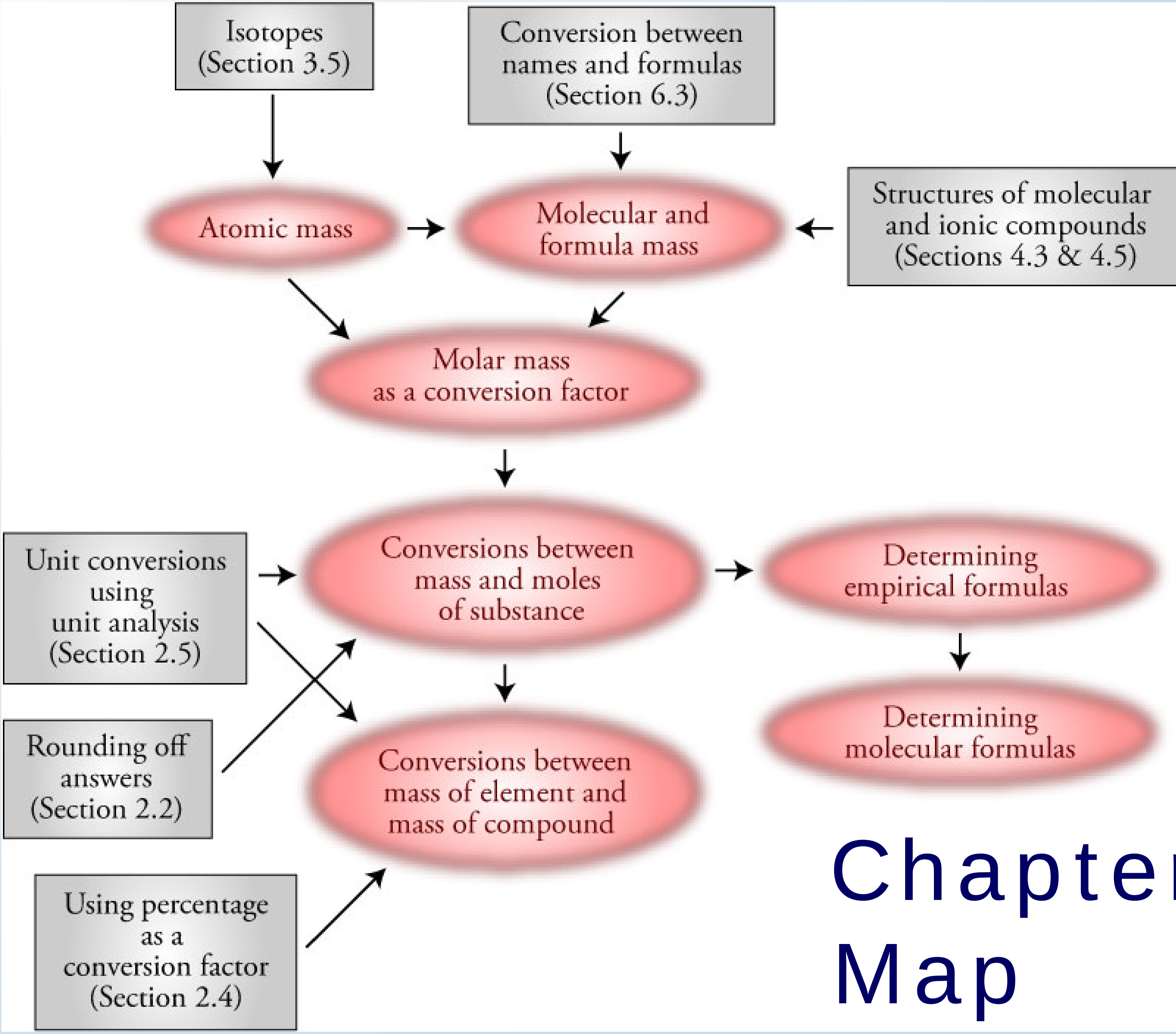

# Chapter 9

## Chemical Calculations and Chemical Formulas

# Chapter Map


Isotopes (Section 3.5)
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Atomic mass
Molecular and formula mass
Structures of molecular and ionic compounds (Sections 4.3 & 4.5)
Molar mass as a conversion factor
Unit conversions using unit analysis (Section 2.5)
Conversions between mass and moles of substance
Determining empirical formulas
Rounding off answers (Section 2.2)
Determining molecular formulas
Conversions between mass of element and mass of compound
Using percentage as a conversion factor (Section 2.4)

# Making Phosphoric Acid

- Furnace Process for making $H_3PO_4$ to be used to make fertilizers, detergents, and pharmaceuticals.
  - React phosphate rock with sand and coke at 2000 °C.

    $$2Ca_3(PO_4)_2 + 6SiO_2 + 10C \rightarrow 4P + 10CO + 6CaSiO_3$$

  - React phosphorus with oxygen to get tetraphosphorus decoxide.

    $$4P + 5O_2 \rightarrow P_4O_{10}$$

  - React tetraphosphorus decoxide with water to make phosphoric acid.

    $$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$$

# Sample Calculations (1)

- What is the maximum mass of $P_4O_{10}$ that can be formed from $1.09 \times 10^4$ kg P?
- The formula for $P_4O_{10}$ provides us with a conversion factor that converts from units of P to units of $P_4O_{10}$.

$$\left(\frac{1 \text{ molecule } P_4O_{10}}{4 \text{ atoms P}}\right)$$

# Sample Calculations (2)

- What is the minimum mass of water that must be added to $2.50 \times 10^4$ kg $P_4O_{10}$ to form phosphoric acid in the following reaction?

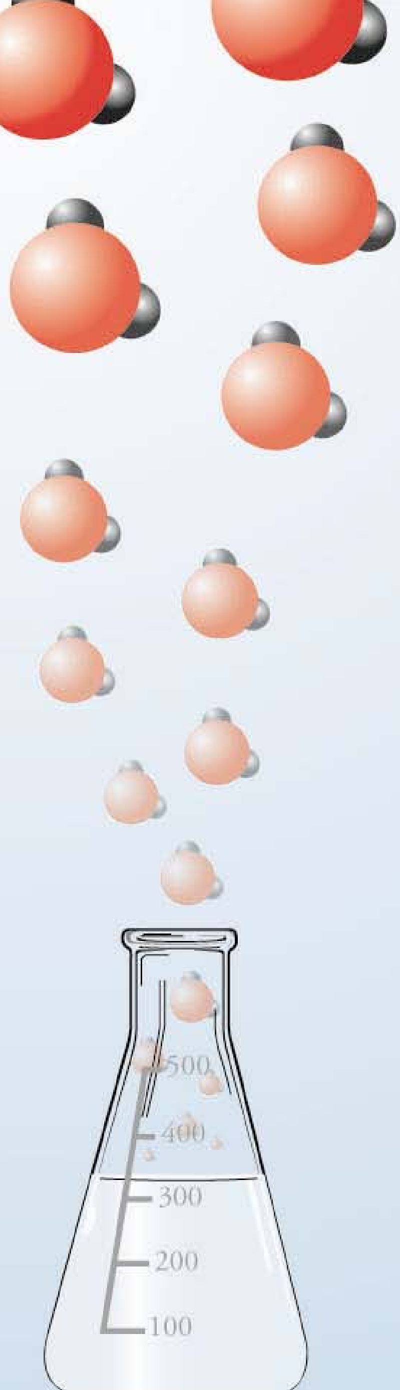
$$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$$

- The coefficients in the balanced equation provide us with a conversion factor that converts from units of $P_4O_{10}$ to units of $H_2O$.

$$\left( \frac{6 \text{ molecules } H_2O}{1 \text{ molecule } P_4O_{10}} \right)$$

# Goal: To develop conversion factors that will convert between a measurable property (mass) and number of particles

| | |
|---|---|
| Measurable Property 1 | Mass 1 |
| ↓ | ↓ |
| Number of Particles 1 | Number of Particles 1 |
| ↓ | ↓ |
| Number of Particles 2 | Number of Particles 2 |
| ↓ | ↓ |
| Measurable Property 2 | Mass 2 |




# Counting by Weighing for Nails

- **Step 1:** Choose an easily measurable property.
  - Mass for nails
- **Step 2:** Choose a convenient unit for measurement.
  - Pounds for nails

# Counting by Weighing for Nails (cont)

- **Step 3:** If the measurable property is mass, determine the mass of the individual objects being measured.
  - Weigh 100 nails: 82 are 3.80 g, 14 are 3.70 g, and 4 are 3.60 g
- **Step 4:** If the objects do not all have the same mass, determine the weighted average mass of the objects.

$$0.82(3.80\text{ g}) + 0.14(3.70\text{ g}) + 0.04(3.60\text{ g}) = 3.78\text{ g}$$

# Counting by Weighing for Nails (cont)

- **Step 5:** Use the conversion factor from the weighted average to make conversions between mass and number of objects.

$$? \text{ nails} = 218 \cancel{\text{lb}} \cancel{\text{nails}} \left(\frac{453.6 \cancel{\text{g}}}{1 \cancel{\text{lb}}}\right)\left(\frac{1 \text{ nail}}{3.78 \cancel{\text{g}} \cancel{\text{nails}}}\right) = 2.62 \times 10^4 \text{ nails}$$

# Counting by Weighing for Nails (cont)

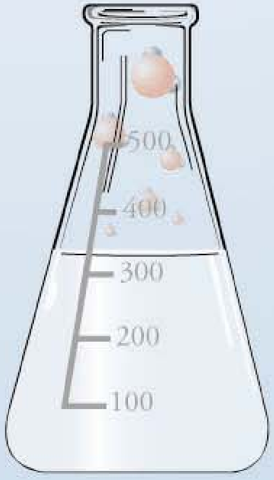
- Step 6: Describe the number of objects in terms of a collective unit such as a dozen, a gross, or a ream.

$$\frac{?\ \text{g nails}}{1\ \text{gross nails}} = \left(\frac{3.78\ \text{g nails}}{1\ \cancel{\text{nail}}}\right)\left(\frac{144\ \cancel{\text{nails}}}{1\ \text{gross nails}}\right) = \frac{544\ \text{g nails}}{1\ \text{gross nails}}$$

$$?\ \text{gross nails} = 218\ \cancel{\text{lb}}\ \cancel{\text{nails}}\left(\frac{453.6\ \cancel{\text{g}}}{1\ \cancel{\text{lb}}}\right)\left(\frac{1\ \text{gross nails}}{544\ \cancel{\text{g}}\ \cancel{\text{nails}}}\right) = 182\ \text{gross nails}$$

# Counting by Weighing for Carbon Atoms

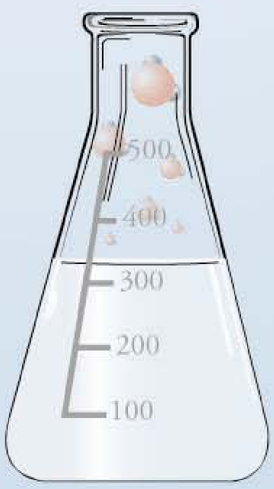
- **Step 1:** Choose an easily measurable property.
  - Mass for carbon atoms
- **Step 2:** Choose a convenient unit for measurement.
  - Atomic mass units (u) for carbon atoms
  - Atomic mass unit (u) = 1/12 the mass of a carbon-12 atom (with 6 p, 6 n, and 6 $e^{-}$)



# Counting by Weighing for Carbon Atoms (cont)

- **Step 3:** If the measurable property is mass, determine the mass of the individual objects being measured.
  - For carbon: 98.90% are 12 u and 1.10% are 13.003355
- **Step 4:** If the objects do not all have the same mass, determine the weighted average mass of the objects.

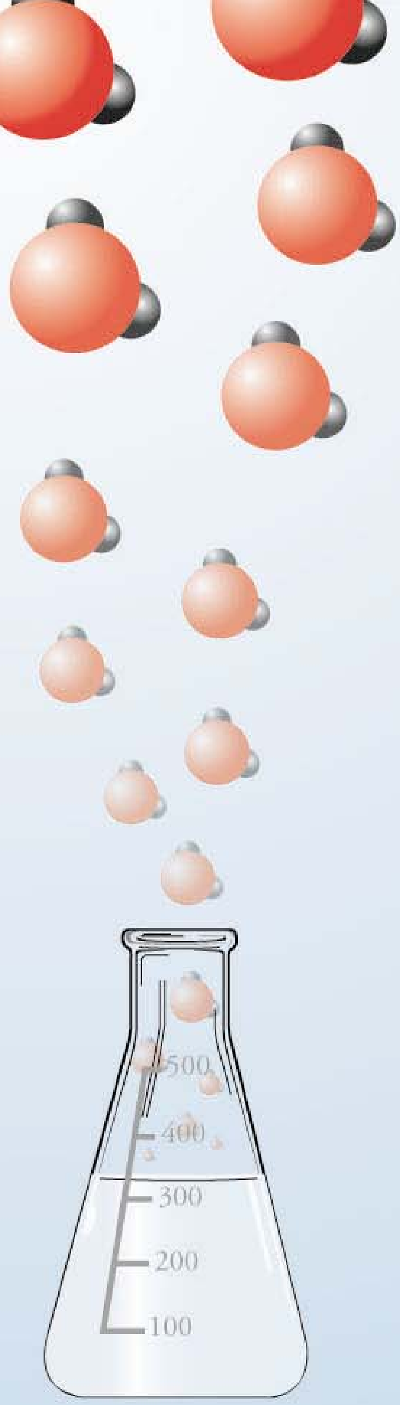
$$0.9890(12\text{ u}) + 0.0110(13.003355\text{ u}) = 12.011\text{ u}$$



# Counting by Weighing for Carbon Atoms (cont)

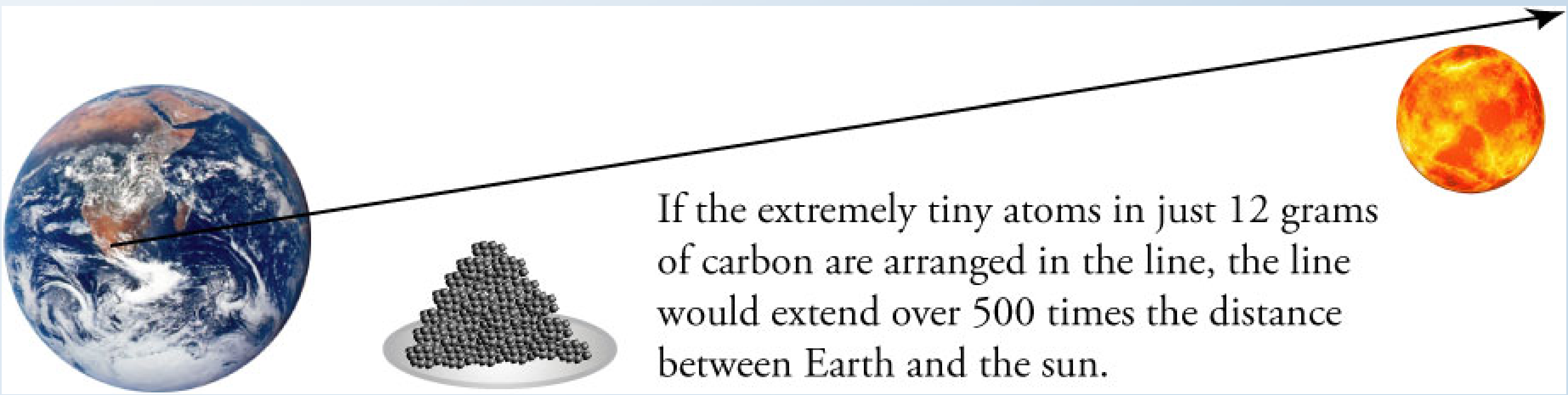
- Step 5: Describe the number of objects in terms of a collective unit such as a dozen, a gross, or a ream, and use this and the weighted average to create a conversion factor to make conversions between mass and number of objects.

# Mole




- A ***mole*** (mol) is an amount of substance that contains the same number of particles as there are atoms in 12 g of carbon-12.
- To four significant figures, there are $6.022 \times 10^{23}$ atoms in 12 g of carbon-12.
- Thus a mole of natural carbon is the amount of carbon that contains $6.022 \times 10^{23}$ carbon atoms.
- The number $6.022 \times 10^{23}$ is often called ***Avogadro's number***.

# Avogadro’s Number



# Molar Mass Development

From the definition of an unified atomic mass unit, u $\frac{12\text{ u C-12}}{1\text{ atom C-12}}$

From the definition of mole $\frac{12\text{ g C-12}}{1\text{ mol C-12}}$

From relative atomic masses

$$\frac{12.011\text{ g C}}{1\text{ mol C}} \quad \frac{24.3050\text{ g Mg}}{1\text{ mol Mg}} \quad \frac{15.9994\text{ g O}}{1\text{ mol O}} \quad \frac{1.00794\text{ g H}}{1\text{ mol H}}$$

# Molar Mass For Elements

- Atomic Mass from the Periodic Table

$$\left(\frac{\text{(atomic mass) g element}}{\text{1mol element}}\right)$$

# Molar Mass Calculation for Carbon

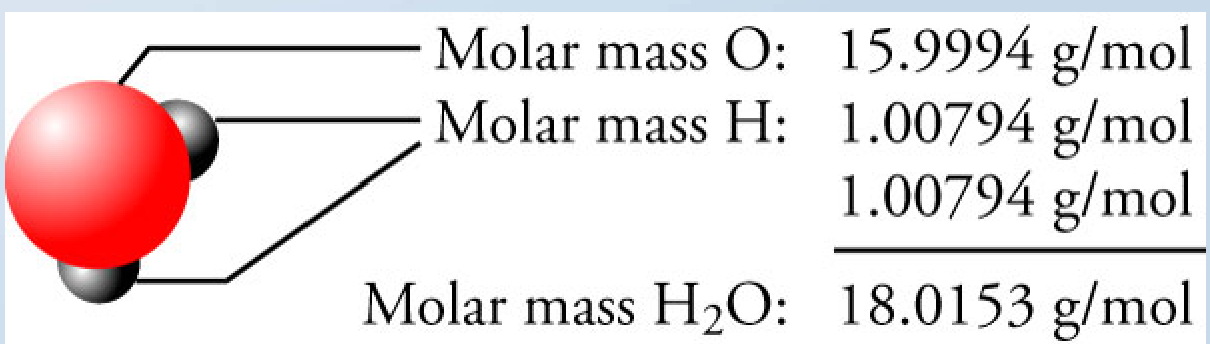
$$? \text{ mol C} = 0.55 \cancel{\text{carat C}} \left(\frac{1 \cancel{\text{g}}}{5 \cancel{\text{carat}}}\right)\left(\frac{1 \text{ mol C}}{12.011 \cancel{\text{g C}}}\right) = 9.2 \times 10^{-3} \text{ mol C}$$

# Goal: To develop conversion factors that will convert between a measurable property (mass) and number of particles

| | |
|---|---|
| Measurable Property 1 | Mass 1 |
| ↓ | ↓ |
| Number of Particles 1 | Moles 1 |
| ↓ | ↓ |
| Number of Particles 2 | Moles 2 |
| ↓ | ↓ |
| Measurable Property 2 | Mass 2 |

# Molecular Mass

- Whole = sum of parts
- mass of a molecule = sum of the masses of the atoms in the molecule
- **molecular mass** = the sum of the atomic masses of the atoms in the molecule



# Molar Mass For Molecular Compounds

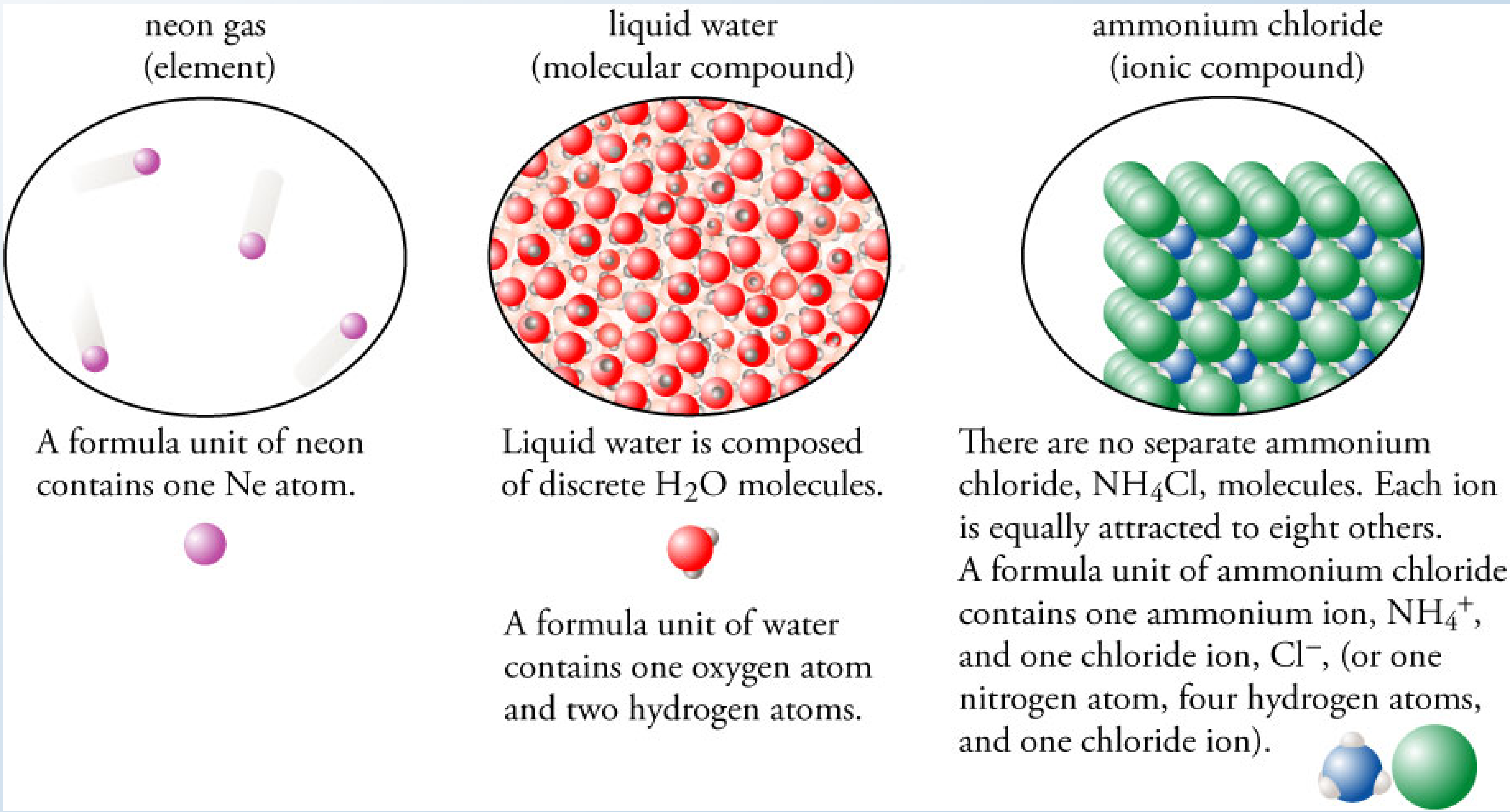
- ***Molecular Mass*** = Sum of the atomic masses of atoms in one molecule

$$\left(\frac{(\text{molecular mass})\ \text{g}\ \text{molecular compound}}{1\ \text{mol}\ \text{molecular compound}}\right)$$

# Formula Units

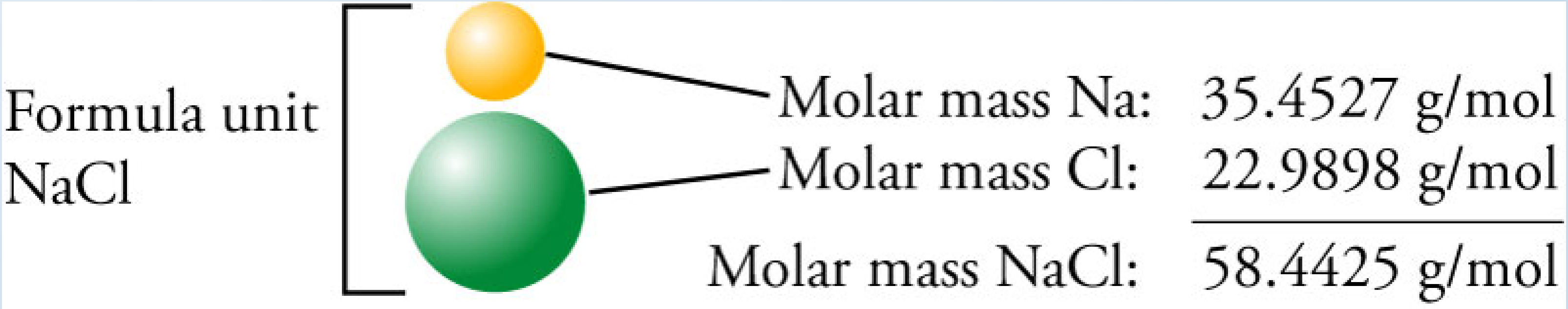
- A ***formula unit*** of a substance is the group represented by the substance’s chemical formula, that is, a group containing the kinds and numbers of atoms or ions listed in the chemical formula.
- Formula unit is a general term that can be used in reference to elements, molecular compounds, or ionic compounds.

# Formula Unit Examples



# Formula Mass for Ionic Compounds

- Whole = sum of parts
- Mass of a formula unit = sum of the masses of the atoms in the formula unit
- ***Formula mass*** = the sum of the atomic masses of the atoms in the formula



# Molar Mass For Ionic Compounds

- ***Formula Mass*** = Sum of the atomic masses of the atoms in a formula unit

$$\left(\frac{(\text{formula mass})\ \text{g ionic compound}}{1\ \text{mol ionic compound}}\right)$$

# Molar Mass Development

From the definition of an unified atomic mass unit, u $\quad \dfrac{12 \text{ u C-12}}{1 \text{ atom C-12}}$

From the definition of mole $\quad \dfrac{12 \text{ g C-12}}{1 \text{ mol C-12}}$

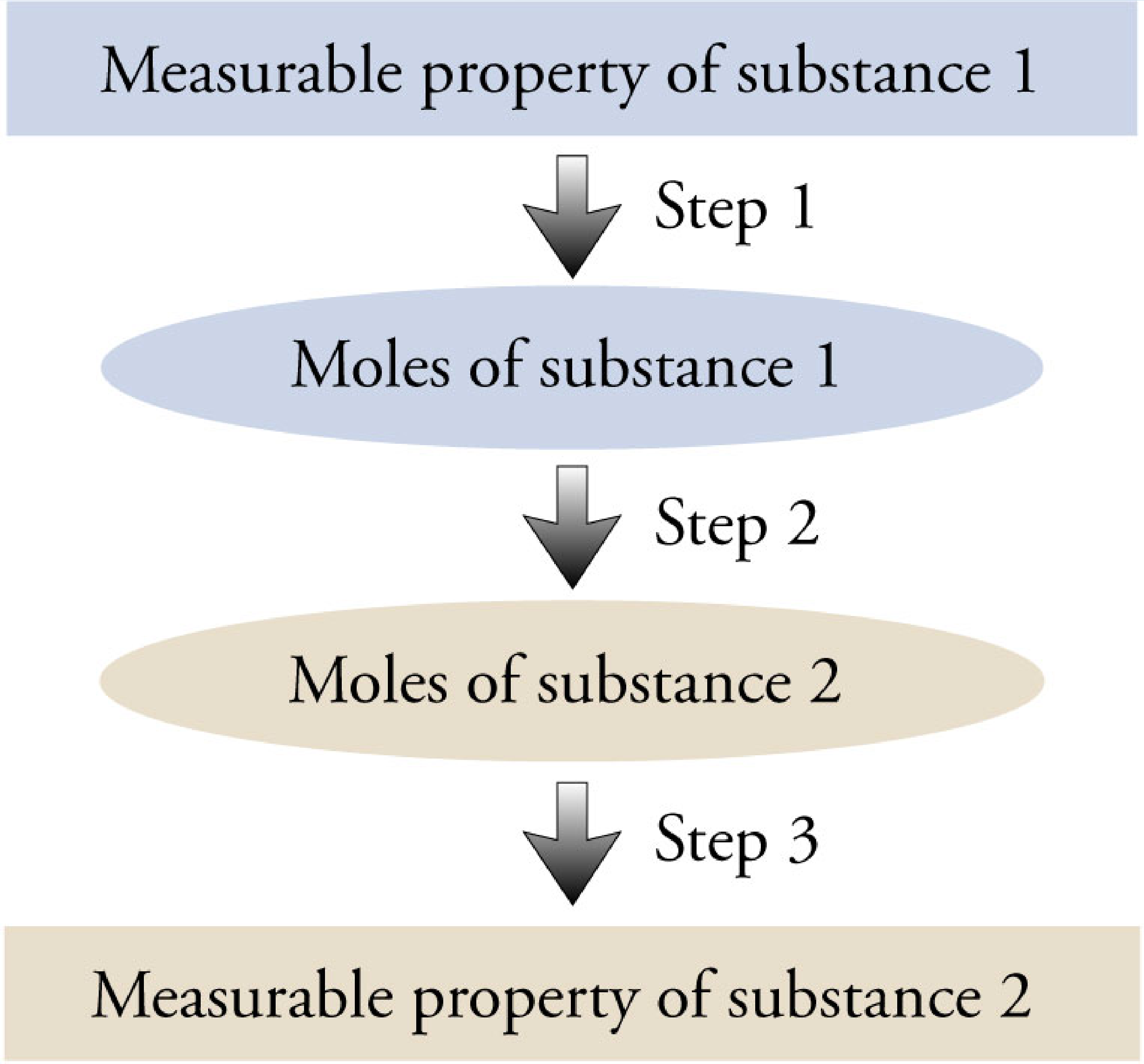
From relative atomic masses

$$\frac{12.011 \text{ g C}}{1 \text{ mol C}} \qquad \frac{24.3050 \text{ g Mg}}{1 \text{ mol Mg}} \qquad \frac{15.9994 \text{ g O}}{1 \text{ mol O}} \qquad \frac{1.00794 \text{ g H}}{1 \text{ mol H}}$$

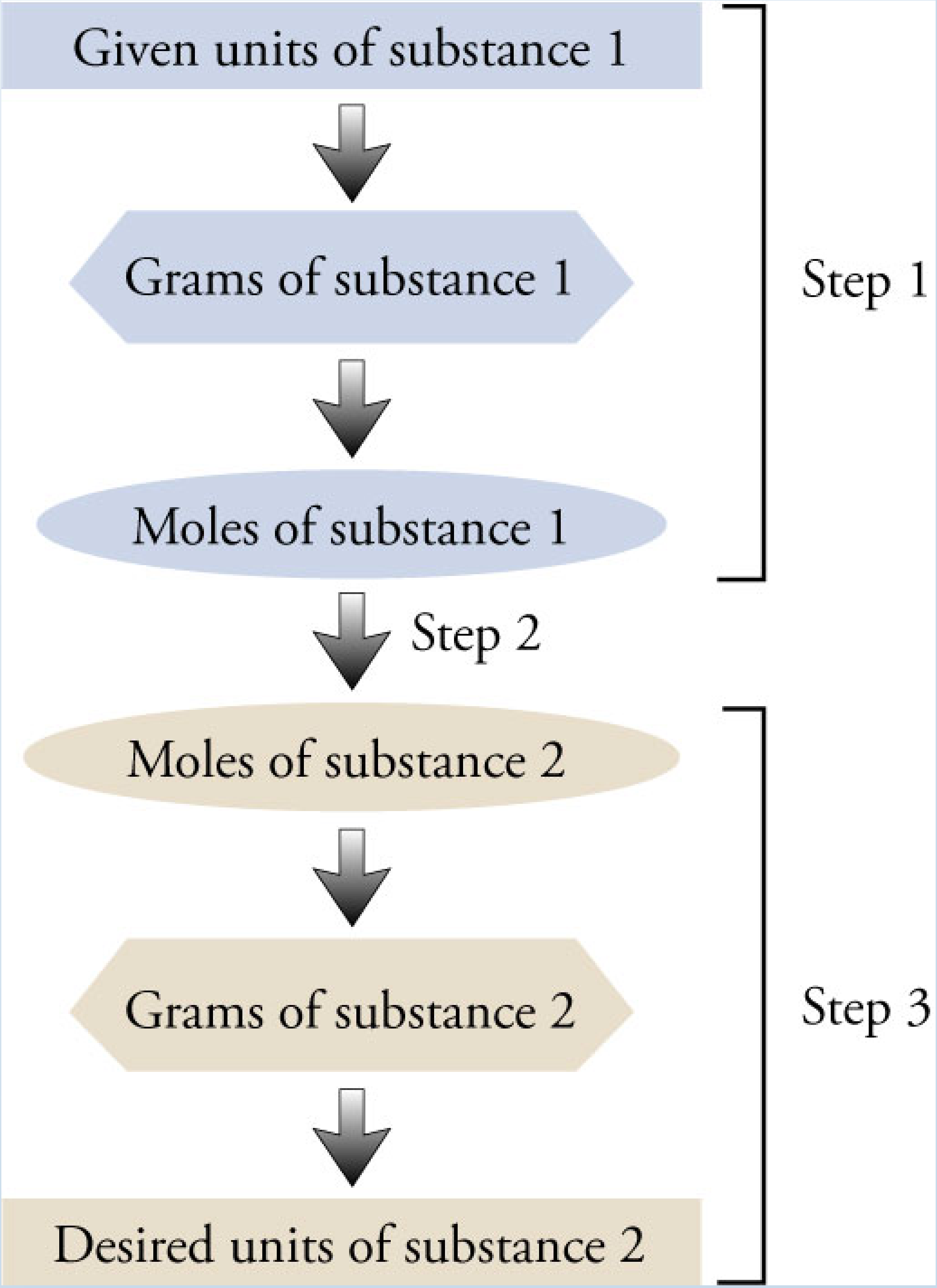
From relative molecular masses $\quad \dfrac{18.0153 \text{ g } H_2O}{1 \text{ mol } H_2O}$

From relative formula masses $\quad \dfrac{58.4425 \text{ g NaCl}}{1 \text{ mol NaCl}}$

# General Conversions



# Units of One Substance to Units of Another



# Study Sheets

- Write a description of the “tip-off” that helps you to recognize the type of problem the calculation represents.
- Write a description of the general procedure involved in the particular type of problem.
- Write an example of the type of calculation.

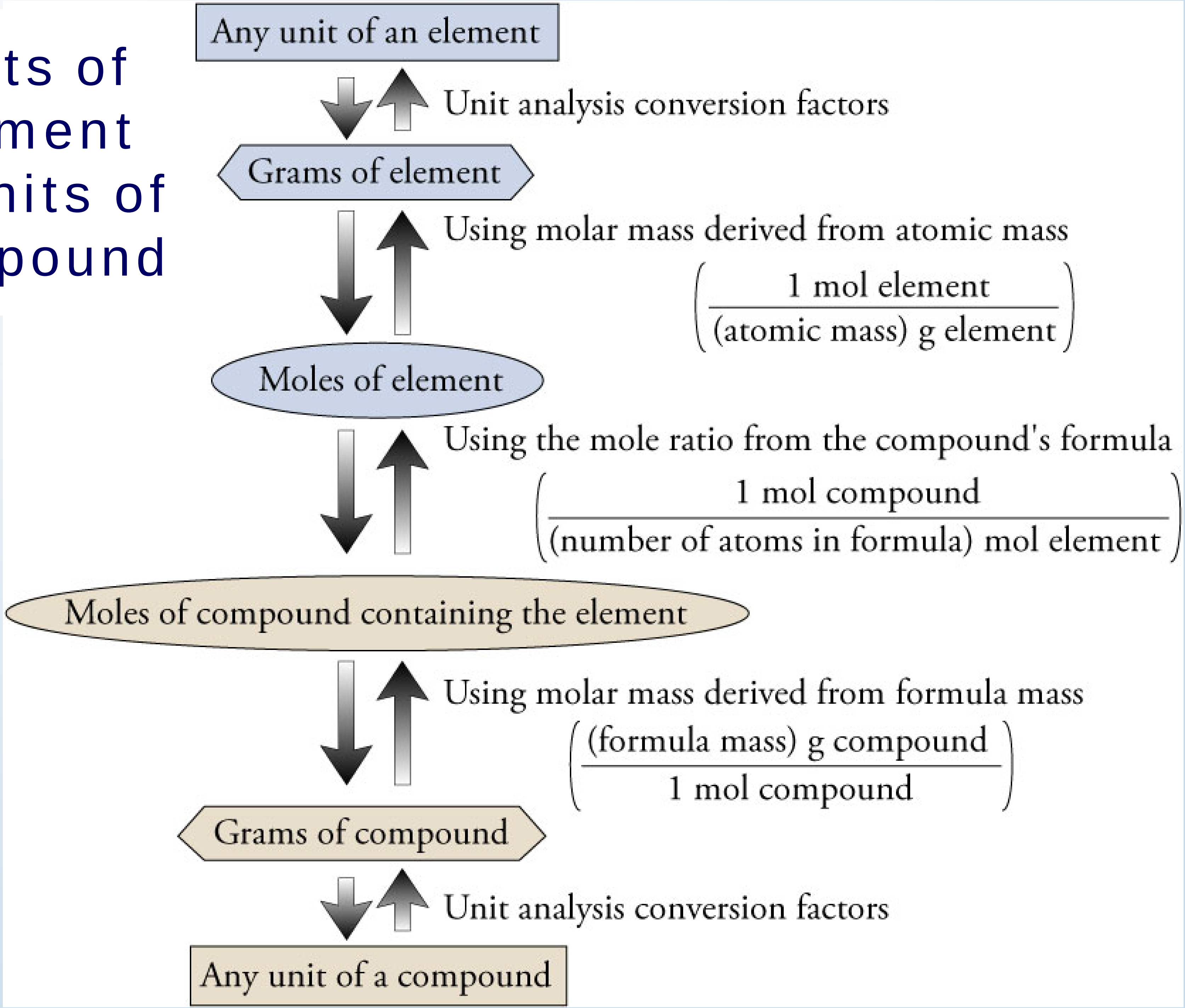
## Sample Study Sheet: Converting Between Mass of Element and Mass of Compound Containing the Element

- **Tip-off:** When you analyze the type of unit you have and the type of unit you want, you recognize that you are converting between a unit associated with an element and a unit associated with a compound containing that element.

# Sample Study Sheet (2)

- **General Steps**
  - – Convert the given unit to moles of the first substance.
  - – Convert moles of the first substance to moles of the second substance using the molar ratio derived from the formula for the compound.
  - – Convert moles of the second substance to the desired units of the second substance.

# Units of Element to Units of Compound



# Calculating Empirical Formulas

**Step 1:** If you are not given mass in grams for each element, convert the data you are given to grams of each element.

- This may involve simple unit conversions. For example, you may be given pounds or milligrams, which you convert to grams using dimensional analysis.
- Sometimes you are given the percentage of each element in the compound. Assume that you have 100 g of compound, and change the units of the values given for the percentages to grams.

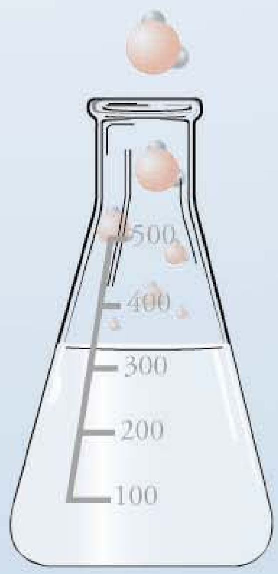
# Calculating Empirical Formulas

**Step 2:** Convert grams of each element to moles by dividing by the atomic mass of the element.

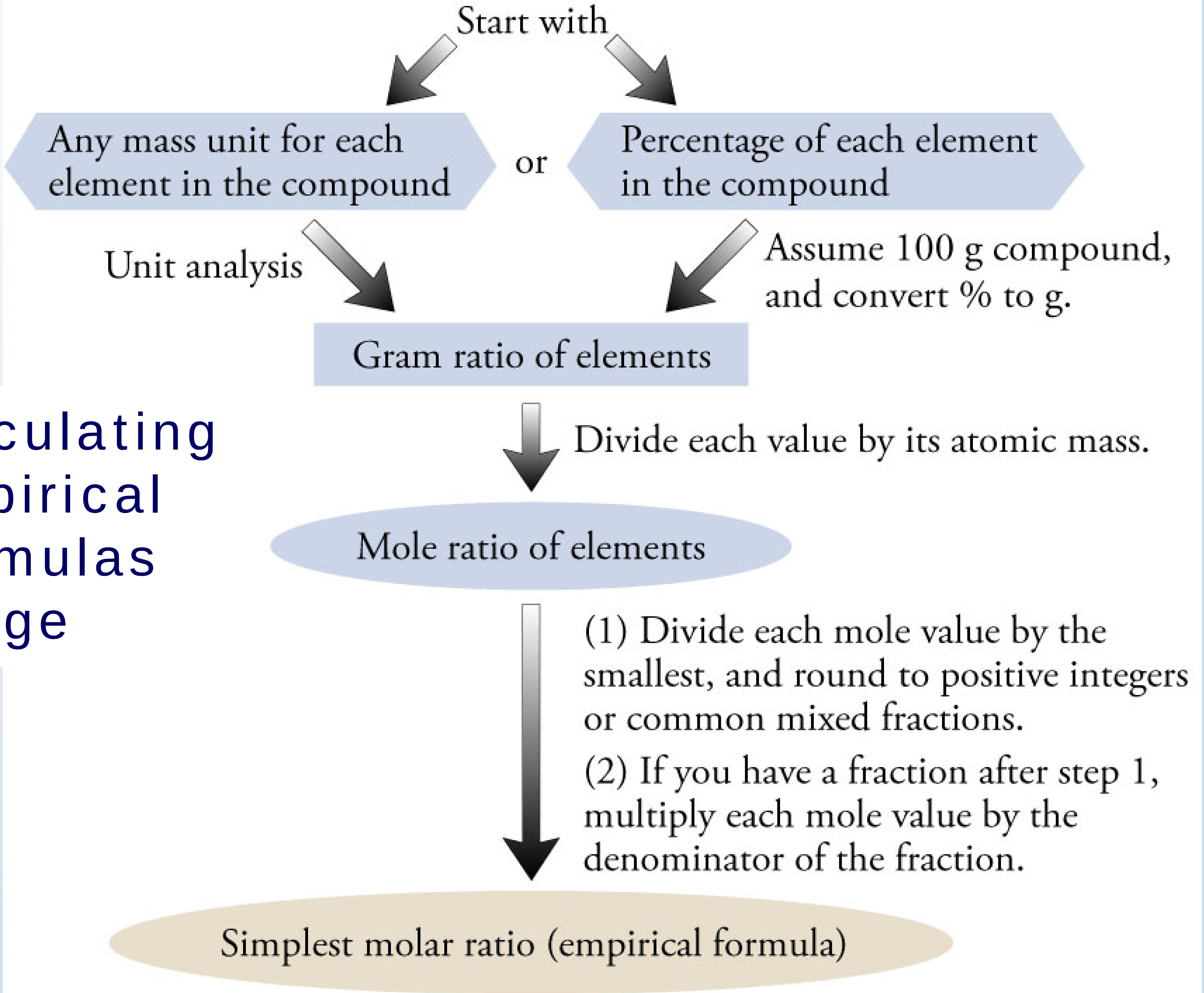
**Step 3:** Divide each mole value by the smallest and round your answers to whole numbers or common mixed fractions.

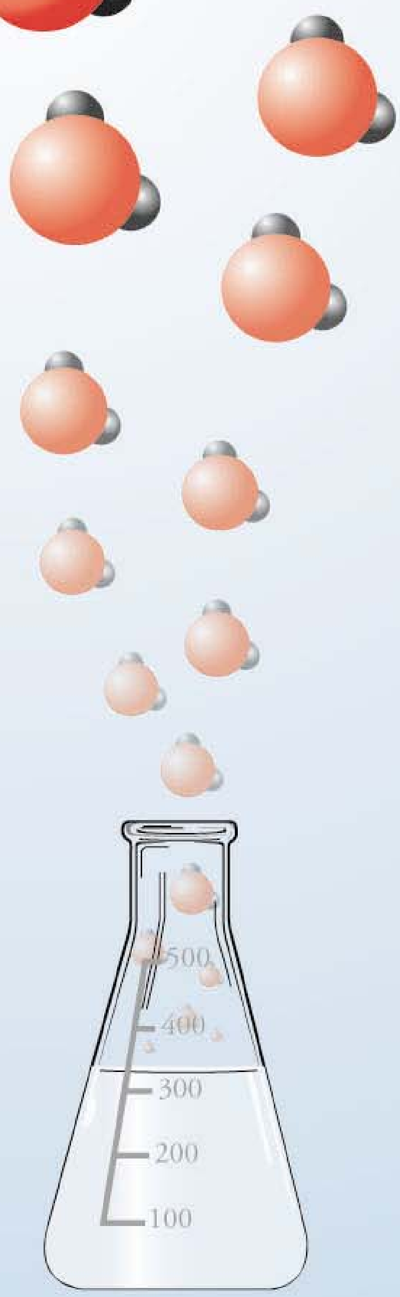
**Step 4:** If you have a fraction after the last step, multiply all the mole values by the denominator of the fraction.

**Step 5:** The resulting mole values correspond to the subscripts in the empirical formula.



# Calculating Empirical Formulas Image






# Calculating Molecular Formulas

**Step 1:** If necessary, calculate the empirical formula of the compound from the data given.

**Step 2:** Divide the given molecular mass by the empirical formula mass.

$$n = \frac{\text{molecular mass}}{\text{empirical formula mass}}$$

**Step 3:** Multiply each of the subscripts in the empirical formula by n to get the molecular formula.

# Calculating Molecular Formulas Image