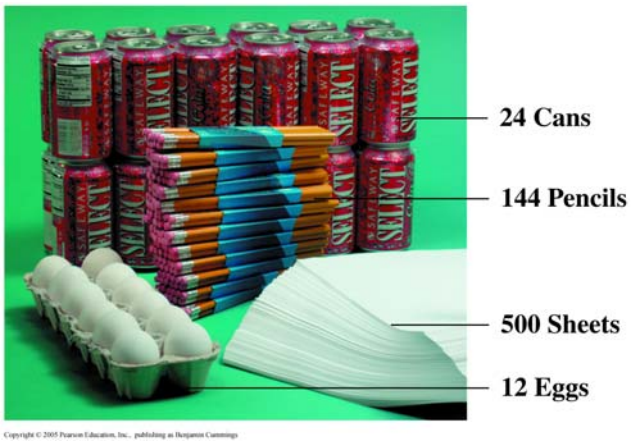
Chapter 6
Chemical Quantities



Atomic Mass

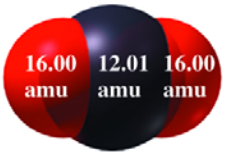
- The atomic mass of any element on the periodic table is expressed in **atomic mass units (amu)**
- This value is listed on the periodic table and represents the mass of a single atom of an element.
- This relationship can be used to write conversion factors
- For example, the atomic mass of iron is 55.85 **amu**, so:

$$\frac{55.85\ \text{amu}}{1\ \text{Fe atom}} \quad \textbf{OR} \quad \frac{1\ \text{Fe atom}}{55.85\ \text{amu}}$$

Formula Mass

- The formula mass of a compound is the sum of the atomic masses of all the atoms in its formula.
- To determine the formula mass, you multiply each element's atomic mass by its formula subscript and then add them all up.

1 Molecule of CO_2



1 C (12.01 amu/C) + 2 O (16.00 amu/O) = 44.01 amu

Avogadro's Number

- How do we keep track of atoms?
  - They are very small, so we group them in a **large** bunch
- We use **Avogadro's Number** to represent this bunch of atoms.
  - Avogadro's Number (symbol *N*) was experimentally determined to be the number of atoms in 12.01 grams of carbon.
- Its numerical value is **6.02×10^{23}**.
  - Therefore, a 12.01 g sample of carbon contains 6.02×10^{23} carbon atoms.

The Mole

- The mole (mol and abbreviated **n**) is a unit of measure for an amount of a chemical substance.
- A mole is Avogadro's number of atoms, that is 6.02×10^{23} atoms.

$1\ mol = 6.02 \times 10^{23}$ atoms

- Notice how this mole relationship resembles one of our unit equations.
  - Therefore, we can use to write conversion factors to convert between the number of atoms and the mass of a substance.

Dimensional Analysis (DA) is Back!

- We will be using the *Dimensional Analysis Method* again that we learned in chapter 3.

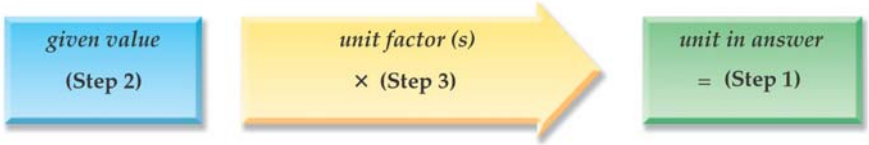
Step 1: Write down the given value related to the answer (What you have)

Step 2: Write down the unit asked for in the answer (What you want)

Step 3: Write out a plan that will let you convert from what you have to what you want.

Step 4: Determine which conversion factor(s) is needed to convert the unit in the given value to the unit in the answer.

Step 5: Check your digits!



Mole Calculations – Moles to Atoms

- How many sodium atoms are in 0.120 mol Na?
  - **Step 1**: Determine what you have: 0.120 mol Na
  - **Step 2**: Determine what you want: ??? atoms of Na
  - **Step 3**: Write a plan to convert from what you have to what you want.
  - **Step 4:** Select conversion factor(s) that allows you to perform your plan

Mole Calculations – Atoms to Moles

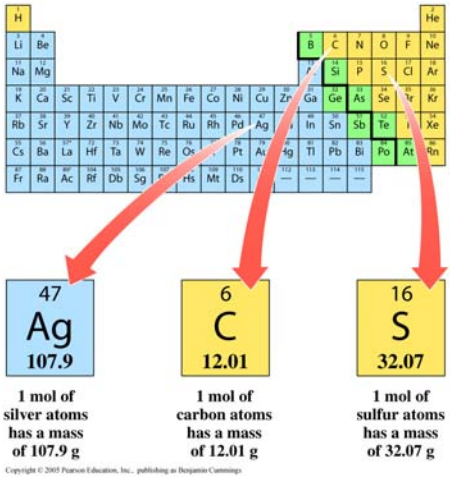
- How many moles of potassium are in 1.25×10^{21} atoms K?
  - **Step 1**: Determine what you have: 1.25×10^{21} atoms K
  - **Step 2**: Determine what you want: ??? moles K
  - **Step 3**: Write a plan to convert from what you have to what you want.
  - **Step 4:** Select conversion factor(s) that allows you to perform your plan

Molar Mass

- The atomic mass of any substance expressed in **grams per mole (g/mol)** is the ***molar mass*** (MM) of that substance.

$$\text{Molar Mass} = \frac{\text{Mass}}{\text{mole}} = \frac{m}{n}$$

- If the atomic mass of iron is 55.85 **amu,** then the molar mass of iron is 55.85 **g/mol**.



Calculating Molar Mass

- The molar mass of a substance is the sum of the molar masses of each element present in the substance.

What is the molar mass of magnesium nitrate?

Mole Calculations – Grams to Moles

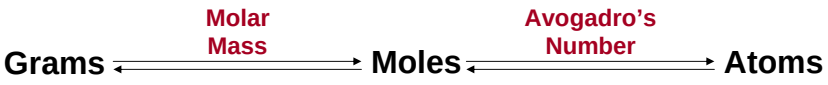
- What is the mass of 1.33 moles of titanium, Ti?
  - **Step 1:** Determine what you have: 1.33 mol Ti
  - **Step 2:** Determine what you want: ??? g Ti
  - **Step 3**: Write a plan to convert from what you have to what you want.
  - **Step 4:** Select conversion factor(s) that allows you to perform your plan

Mole Calculations – Grams to Atoms

- Now we will use the molar mass of a compound to convert between grams of a substance and moles or particles of a substance.

6.02×10^{23} particles = **1 mol** = molar mass

- If we want to convert particles to mass, we must first convert particles to moles and then we can convert moles to mass.



Mole Calculations – Atoms to Grams

- What is the mass of 2.55×10^{23} atoms of lead?
  - **Step 1:** Determine what you have: 2.55×10^{23} atoms of lead
  - **Step 2:** Determine what you want: ??? g Pb
  - **Step 3**: Write a plan to convert from what you have to what you want.
  - **Step 4:** Select conversion factor(s) that allows you to perform your plan

Mole Calculations – Grams to Atoms

- How many O_2 molecules are present in 0.470 g of oxygen gas?
  - **Step 1:** Determine what you have: 0.470 g O_2
  - **Step 2:** Determine what you want: ??? Molecules of O_2
  - **Step 3**: Write a plan to convert from what you have to what you want.
  - **Step 4:** Select conversion factor(s) that allows you to perform your plan

The Percent Concept

- A ***percent***, %, expresses the amount of a single portion compared to an entire sample.

$$\% = \frac{\text{portion of interest ("Part")}}{\text{total sample ("Whole")}} \times 100\%$$

- The ***percent composition*** of a compound lists the mass percent of each element.0

$$\% = \frac{\text{Mass of element}}{\text{Mass of compound}} \times 100\%$$

Calculating Percent Composition

- There are a few steps to calculating the percent composition of a compound. Lets practice using H_2O.

Step 1: Assume you have 1 mole of the compound.

Step 2: Determine the moles of each element in one mole of H_2O

- H_2O contains 2 mol of hydrogen and 1 mol of oxygen.

Step 3: Calculate the Molar Mass

$2(1.01\text{ g H}) + 1(16.00\text{ g O}) = \text{molar mass } H_2O$

$2.02\text{ g H} + 16.00\text{ g O} = \mathbf{18.02\text{ g } H_2O}$

Calculating Percent Composition

Step 4: Next, find the percent composition of water by dividing the mass of the part (the element) by the mass of the whole (Water) then multiplying by 100:

$$\frac{2.02 \text{ g H}}{18.02 \text{ g } H_2O} \times 100\% = \mathbf{11.2\% \text{ H}}$$

$$\frac{16.00 \text{ g O}}{18.02 \text{ g } H_2O} \times 100\% = \mathbf{88.79\% \text{ O}}$$

Empirical Formulas

- The ***empirical formula*** of a compound is the **simplest whole number ratio** of elements in a formula unit
- The ***molecular formula*** of a compound is some multiple of the empirical formula
  - However, the molecular formula and empirical formula can be the same for a compound too!
- **For example**:
  - The molecular formula of benzene is C_6H_6. What is its empirical formula?
  - The molecular formula of octane is C_8H_{18}. What is its empirical formula?

Finding Empirical Formulas

- There are 4 to 5 steps to determining the empirical formula of a compound:
  - **Step 1:** Determine the grams of each element in the formula.
    - Can be given in problem as actual grams or as percent composition
  - **Step 2:** Convert the grams of each element into moles of that element with its molar mass.
  - **Step 3:** Determine which element has the smallest number of moles and divide **ALL** the elemental moles by this number to get whole numbers.
  - **Step 4**: If you do not get all whole numbers with step 3, multiply all the values from step 3 by the same whole number factor to get whole numbers.
  - **Step 5:** These whole numbers are the subscripts for the elements in the compounds so write the correct formula.

Percent to Mass; Mass to Moles; Divide by Smallest; Multiply till Whole!

Calculating Empirical Formulas

- A 1.640 g sample of radium metal was heated to produce 1.755 g of radium oxide. What is the empirical formula?

Calculating Empirical Formulas

- Benzene is 92.2% carbon and 7.83% hydrogen. What is the empirical formula?

Converting Decimals to Whole Numbers

- When calculating empirical formulas, you don't always get a nice whole number.
- Sometimes the result of dividing by the smallest number of moles gives a decimal instead.
- Decimal values that are close to a whole number can be rounded to that number:

2.04 becomes 2.00 and 6.98 becomes 7.00

- However, a decimal that is greater than 0.1 or less than 0.9 has to be multiplied by a small integer:

Table 7.4 Some Multipliers That Convert Decimals to Whole Numbers

Decimal	Multiply by	Example		Whole number
0.20	5	1.20 × 5	=	6
0.25	4	2.25 × 4	=	9
0.33	3	1.33 × 3	=	4
0.50	2	2.50 × 2	=	5
0.67	3	1.67 × 3	=	5

Molecular Formulas

- The empirical formula for benzene is CH. This represents the ratio of C to H atoms in a benzene molecule.
- The actual ***molecular formula*** is some multiple (***f***) of the empirical formula, $(CH)_f$.
- To determine ***f***, we divide the mass of the molecular formula by the mass of the empirical formula:

$$\frac{\text{Mass of Molecular Formula}}{\text{Mass of Empirical Formula}} = f$$

Finding Molecular Formulas

- There are 4 steps to determining the molecular formula of a compound:
  - **Step 1:** Determine the empirical formula of the compound.
    - May need to calculate yourself or may be given in the problem
  - **Step 2:** Calculate the mass of the empirical formula.
  - **Step 3:** Divide the mass of the molecular formula (usually given in the problem!) by the mass of the empirical formula to determine the multiplier factor (*f*).
  - **Step 4**: Multiply all the subscripts in the empirical formula by the factor *f* to get the molecular formula.

Molecular Formulas

- Benzene has a molar mass of 78.11 g/mol. The empirical formula is CH. Find *n* to then find the molecular formula
  - **Step 1:** Determine the empirical formula. Here its given in the problem: **CH**
  - **Step 2:** Determine the mass of the empirical formula

    CH: 13.02 g/mol
  - **Step 3:** Divide the mass of the molecular formula (given above: 78.11 g/mol) by the mass of the empirical formula to determine ***f***
  - **Step 4:** Multiply all the subscripts in the empirical formula by ***f***

$$\frac{(CH)_f}{CH} = \frac{78.11\ g/mol}{13.02\ g/mol}$$

f = 6 and the molecular formula is $\mathbf{C_6H_6}$