Module 5: Counting Molecules and Atoms in Chemical Equations

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1 Overview

1. **Decomposition Reactions**
2. **Formation Reactions**
3. **Complete Combustion Reactions**
4. **Incomplete Combustion Reactions**
5. **Molecular Mass**
6. A **Mole** is a fixed number of objects (usually molecules or atoms).

2 Classifying Chemical Reactions

2.1 Decomposition Reactions

Definition 1 (Decomposition Reaction) *A reaction that changes a compound into its constituent elements*

- these are reasonably easy to predict
- they often (not always) involve some sort of energy input[1]
- we can think of our old friend: $2\,H_2O\,(l) \longrightarrow 2\,H_2\,(g) + O_2\,(g)$

[1] A stable compound represents a *localized* low-energy state, so getting it to decompose can require some energy input.

2.2 Formation Reactions

Definition 2 (Formation Reaction) *A reaction that starts with two or more elements and produces one compound*

- these are really the opposite of decomposition reactions
- we can even write them by writing decomposition reactions backwards: $2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O\,(l)$

2.3 Complete Combustion Reactions

Definition 3 (Complete Combustion Reaction) *A reaction in which O_2 is added to a compound containing carbon (C) and hydrogen (H), producing CO_2 and H_2O*

- this is a narrow definition of *burning*[2]
- this definition excludes some really exciting combustion reactions, like $2\,Mg\,(s) + O_2\,(g) \longrightarrow 2\,MgO\,(s)$
- technically, our *formation* example is a combustion, but we're classifying that differently for now

2.4 Incomplete Combustion Reactions

- these are combustion reactions where insufficient O_2 means the combustion produces CO or even just C instead of CO_2
- these can be pretty harmful — even dangerous (breathing CO can actually kill you) — but they can also be beneficial[3]
- the main take-away here is just that a combustion reaction can go one of three ways, depending on the O_2 levels in the environment

3 Atomic Mass

We've already talked about The Law of Definite Proportions, and how that led Dalton to formulate theories about atoms. Now we're going to dig more into that...

- every element on the Periodic Table contains two important numbers: the *atomic number* (above the element symbol), and the *atomic mass* (below the symbol)
- remember that an element is made up of only one type of atom, so we'll use the terms "atom" and "element" interchangeably here

[2]This is really a glimpse into the exciting world of *organic chemistry*.
[3]This is, after all, how we make charcoal.

- the atomic number tells us the number of protons (and hence the number of electrons) in the atom[4]
- the atomic mass tells us how much mass the atom has, in atomic mass units (**amu**)
- the relationship between amu and grams is: $1g = 6.02214076 \times 10^{23} amu$
- so one H atom has the mass of $1.00794amu$, the mass of $6.02214076 \times 10^{23}$ H atoms is $1.00794g$
- that means our Periodic Table can help us calculate how many atoms of any given element we have, if we know the mass of the sample:

Example 1 How many atoms are in $1.00kg$ of gold (Au)?

$$\boxed{\begin{aligned}
m_{unit} &= \frac{196.966569amu}{1atom} \\
m_{sample} &= 1.00kg \\
&= (1.00kg)(\frac{1000g}{1kg})(\frac{6.02214076 \times 10^{23}amu}{1g}) \\
&= (1.00\cancel{kg})(\frac{1000\cancel{g}}{1\cancel{kg}})(\frac{6.02214076 \times 10^{23}amu}{1\cancel{g}}) \\
&= 6.02214076 \times 10^{26}amu \\
count &= \frac{m}{m_{unit}} \\
&= (6.02214076 \times 10^{26}amu)(\frac{1atom}{196.966569amu}) \\
&= (6.02214076 \times 10^{26}\cancel{amu})(\frac{1atom}{196.966569\cancel{amu}}) \\
&= 3.0574430933 \times 10^{24}atoms \\
&= 3.05744309 \times 10^{24}atoms
\end{aligned}} \tag{1}$$

4 Molecular Mass

5 Mole

Definition 4 (Mole) *A mole is* $6.02214076 \times 10^{23}$ *objects.*

- a *mole* is a number of objects
- a mole is like a pair, or a dozen, or a gross (see Table 1)

[4]More on this later

- an element's mass in *amu* is the same as a mole of that element's mass in g[5]
- because atoms and molecules are tiny[6], we find it easier to measure out *moles* than *atoms*
- let's consider our old friend: $2\,H_2 + O_2 \longrightarrow 2\,H_2O$
  - so it takes two H_2 molecules and one O_2 molecule to make two H_2O molecules
  - or, we could say it takes two dozen H_2 molecules and one dozen O_2 molecule to make two dozen H_2O molecules
  - or, we could say it takes two gross of H_2 molecules and one gross of O_2 molecule to make two gross of H_2O molecules
  - or, we could say it takes two moles of H_2 molecules and one mole of O_2 molecule to make two mole of H_2O molecules
- technically, chemical reactions occur "per each," but we find it much easier to measure them "per mole"
- so all our chemistry is going to be done "per mole"

6 Homework

Review Problems: p. 161 # 1–10 (not to be turned in)
Practice Problems: p. 162 # 1–10 (due 2025-11-07)
Experiment 5.1, p. 149 (due 2025-11-07)

References

[Wile, 2003] Wile, J. L. (2003). *Exploring Creation with Chemistry.* Apologia Educational Ministries, Inc., 2 edition.

[5] This is why there aren't mass units in the Periodic Table: a single H atom has a mass of $1.01amu$, a mole of H atoms has a mass of $1.01g$.

[6] Like really, really small.

Name	Number
pair	2
trio	3
half-dozen	6
dozen	12
baker's dozen	13
score	20
gross	144
mole	6.022×10^{23}

Table 1: Names of collections of objects