Module 2: What is a "Chemical Species"?

Atomic Theory and Symbols

Fundamentals of Chemistry Open Course

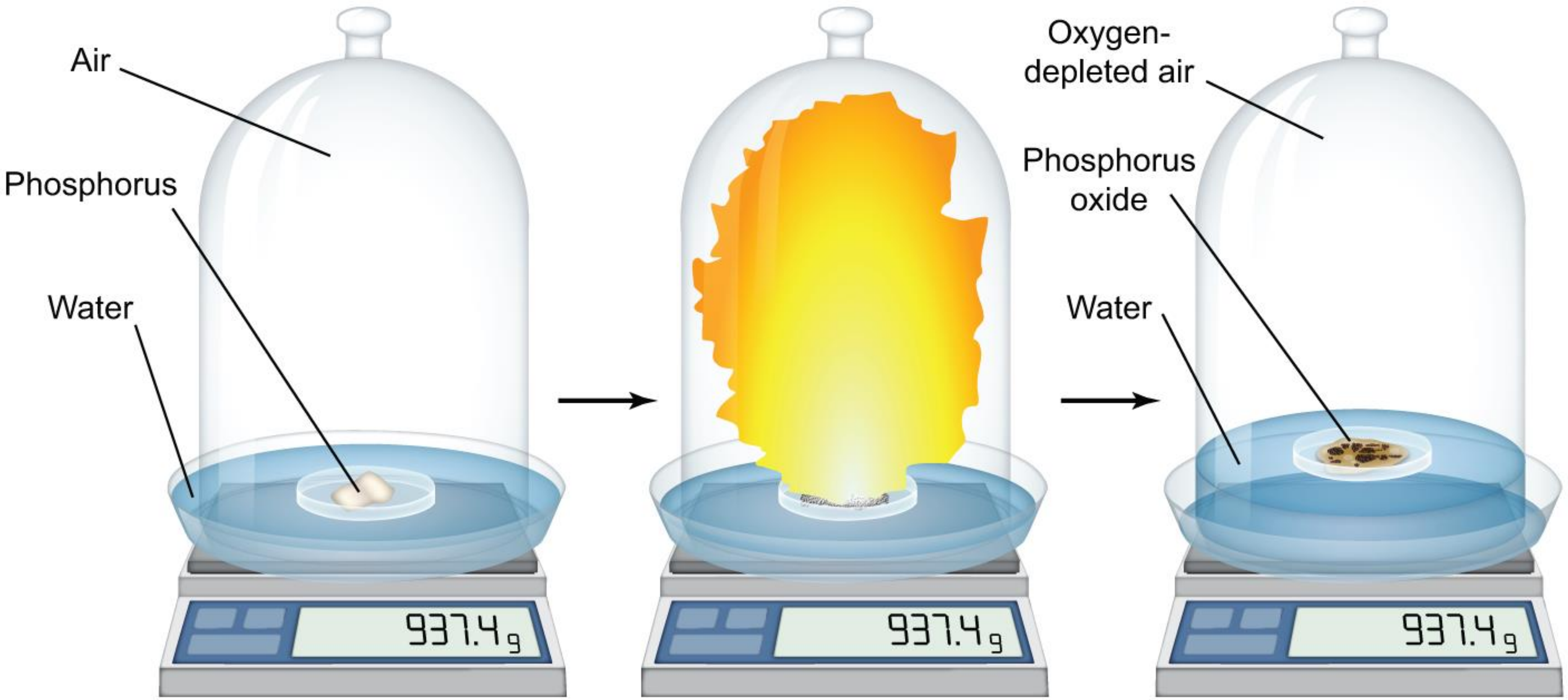
Learning Objectives | Module 2

1. State and apply the laws of chemical combination.
2. State and apply the tenets of the modern atomic theory.
3. Visualize the subatomic particles that constitute the atom using a simple planetary model; count subatomic particles using atomic number (Z) and mass number (M).
4. Represent an atom or ion using an atomic symbol.
5. Represent the number ratios of atoms in a compound using a chemical formula.
6. Visualize and distinguish between submicroscopic models of molecular and ionic compounds.
7. Use the periodic table to efficiently find information about a chemical element.
8. Recognize key collections of elements on the periodic table.
9. Determine the name of a binary ionic compound from the chemical formula and *vice versa*.
10. Determine the name of a simple molecular compound from the chemical formula and *vice versa*.

- **The law of conservation of mass:** the total mass of all matter is constant during a chemical reaction; atoms are neither created nor destroyed during chemical reactions.

Example. When phosphorus is burned in air in a closed container, the total mass of material remains constant.

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



- **The law of definite proportions:** all samples of a compound, regardless of the source or size, contain the same proportions of the elements that constitute the compound.

Example. Potassium carbonate contains potassium, carbon, and oxygen in a ratio of 56.6 : 8.7 : 34.7 by mass.

- **The law of multiple proportions:** when elements combine in multiple proportions to form different compounds, the different proportions are related by whole-number ratios.

Example. There are three oxides of sulfur. The mass ratios in these compounds are related by whole-number factors.

Oxide A contains sulfur and oxygen in a 2 : 1 ratio by mass.

Oxide B contains sulfur and oxygen in a 1 : 1 ratio by mass.

Oxide C contains sulfur and oxygen in a 1 : 2 ratio by mass.

Atomic Theory

- The laws of chemical combination and other observations gave rise to the **atomic theory**.

- **Tenets of the atomic theory:**

  1. Matter is composed of exceedingly small particles called atoms. An atom is the smallest unit of an element that can participate in a chemical change.

  2. An element consists of only one type of atom, which has a mass that is characteristic of the element and is the same for all atoms of that element. A macroscopic sample of an element contains an incredibly large number of atoms, all of which have identical chemical properties.

  3. Atoms of one element differ in properties from atoms of all other elements.

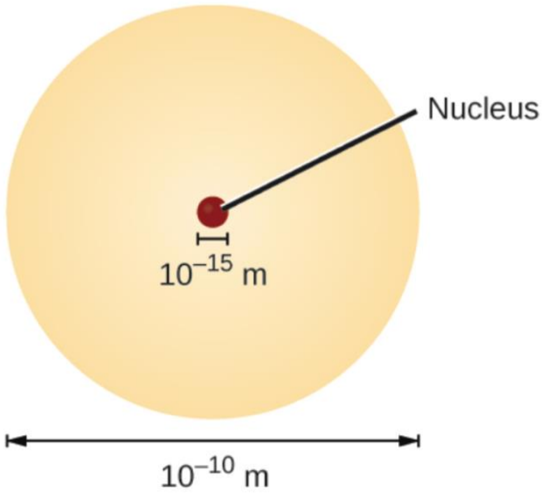
  4. A compound consists of atoms of two or more elements combined in small whole-number ratios.

  5. Atoms are neither created nor destroyed during a chemical change but are instead rearranged to yield substances that are different from those present before the change.

- The atom contains three types of **subatomic particles:**
  - **Protons** are positively charged and sit at the center of the atom in the **nucleus**. **Neutrons** also reside in the nucleus but have neutral charge.
  - **Electrons** are negatively charged, are much lighter than protons and neutrons, and surround the nucleus.

Particle	Charge (*e*)	Mass (g)	Mass (u)	Location
Proton	+1	1.67×10^{-24}	1.0073	In the nucleus
Neutron	0	1.67×10^{-24}	1.0073	In the nucleus
Electron	–1	9.11×10^{-31}	5.49×10^{-4}	Outside the nucleus

The elementary charge *e* is 1.6×10^{-19} Coulombs. One atomic mass unit (u or amu) is 1.67×10^{-24} grams.



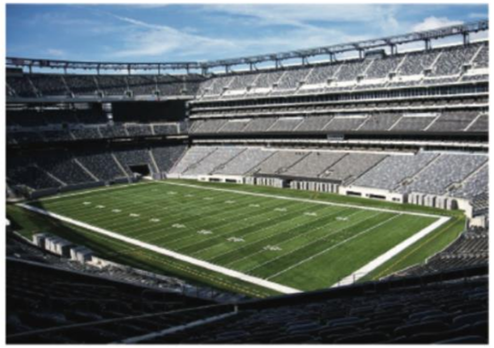


Figure. If the atom were blown up to be as large as a football stadium, the nucleus would be the size of a blueberry.

Atomic Number and Mass Number: Counting Subatomic Particles

- **Atomic number (*Z*):** the number of protons in the nucleus *or* the number of electrons in the neutral atom
- **Mass number (*M*):** the number of protons plus the number of neutrons in the nucleus
- Atomic number *defines* the elements. All atoms of a given element have the same number of protons in the nucleus. Atoms with different numbers of protons in their nuclei correspond to different elements.
- Atoms of a given element may have equal numbers of protons (equal *Z*) but different numbers of neutrons (unequal *M*). Such atoms are called **isotopes**.

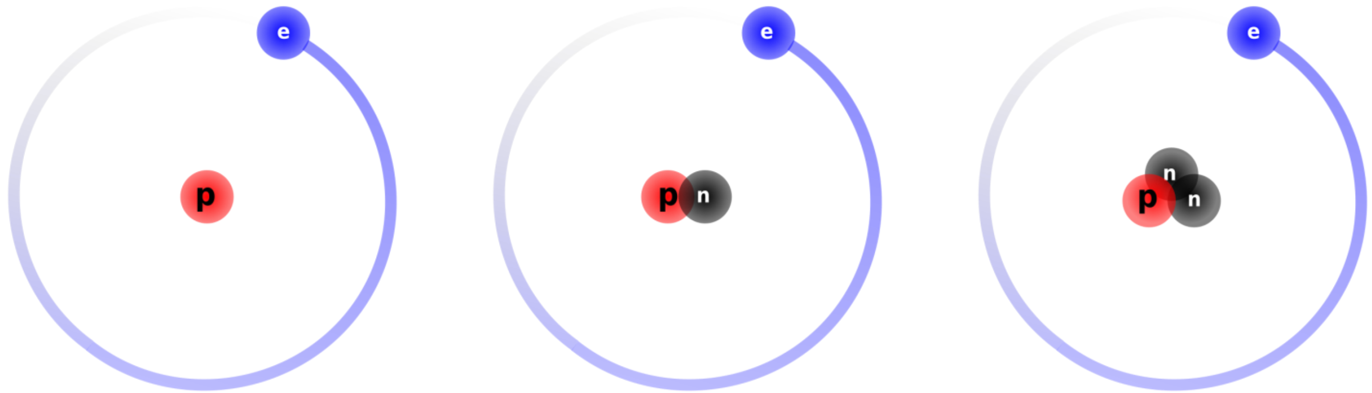


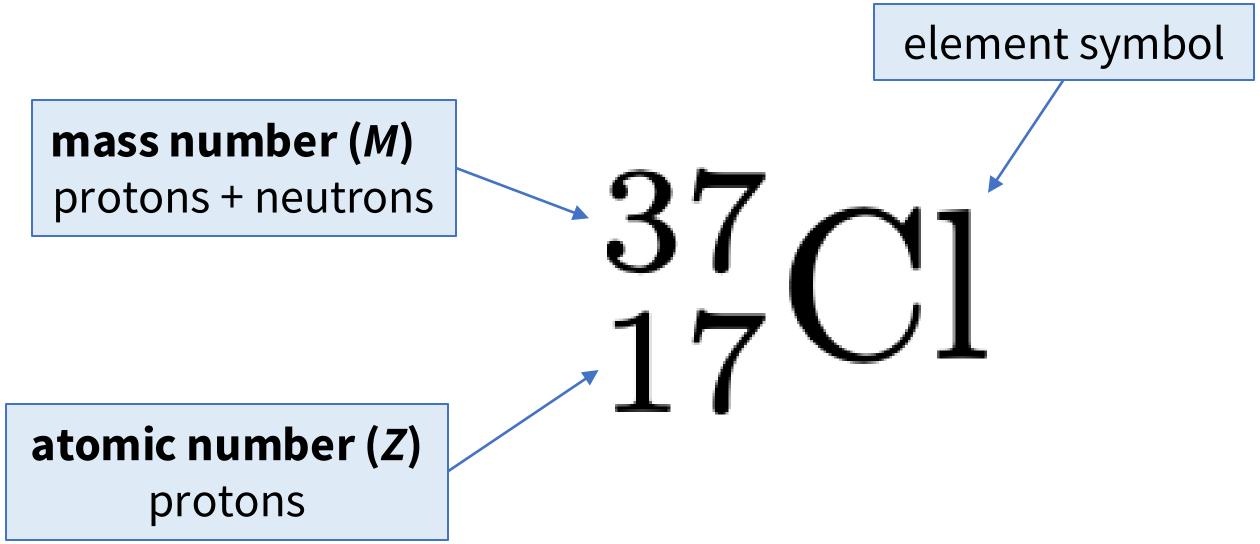
Figure. The three isotopes of hydrogen have equal atomic numbers but unequal mass numbers.

Atomic Symbols

- Each element has an **element symbol** of one or two letters used to signify atoms of the element or a bulk sample of the element.

- In an **atomic symbol**, superscripts and subscripts to the left of the element symbol are used to specify the mass number and atomic number (redundant with the element symbol).

- The number of neutrons is equal to the difference between the mass number and atomic number.



Representing Ions

- Charged particles containing different numbers of protons and electrons are called **ions**.
  - **Cations** contain fewer electrons than protons and net positive charge.
  - **Anions** contain more electrons than protons and net positive charge.

- Ions may consist of a single atom (**monatomic ions**) or multiple atoms bonded together (**polyatomic ions**).

- Charge is indicated with a superscript to the right of the element symbol or formula: 3–, 2–, –, +, 2+, 3+, etc.

charge
30 protons (Zn) and 28 electrons (2+)

$$^{64}\mathrm{Zn}^{2+}$$

charge
35 protons (Br) and 36 electrons (–)

$$^{81}\mathrm{Br}^{-}$$

Lingering Questions

- How do we think about atomic mass in a macroscopic sample of an element with multiple isotopes?
- Do the charges of monatomic ions formed by the elements follow any sort of pattern?
- How similar are the properties of isotopes? Do they display the same chemistry?
- Do electrons really “orbit” the nucleus like planets around the sun? What does an electron really “look like”?