

THE MASS LAWS

- The development and acceptance of atomic theory was based on 3 important scientific laws. These were: the law of conservation of mass, the law of definite proportion, and the law of multiple proportion.
- The *law of conservation of mass* states that in a chemical reaction, matter is neither created or destroyed.
- The *law of definite proportions* states that all samples of a given compound, regardless of their source or preparation, have the same proportions of their constituent elements.
- The *law of multiple proportions* states that when two elements combine to form two different compounds, the atoms combine in simple, whole-number ratios.

Examples:

1. An automobile gasoline tank holds 21 kg of gasoline. When the gasoline burns, 84 kg of oxygen is consumed, and carbon dioxide and water are produced. What is the total combined mass of carbon dioxide and water that is produced?

2. Upon decomposition, one sample of magnesium fluoride produced 1.65 kg of magnesium and 2.57 kg of fluorine. A second sample produced 1.32 kg of magnesium. How many grams of fluorine did the second sample produce?

STRUCTURE OF THE ATOM

- All atoms are composed of the same subatomic particles: protons, neutrons and electrons.

TABLE 2.1 Subatomic Particles

	Mass (kg)	Mass (amu)	Charge (relative)	Charge (C)
Proton	1.67262×10^{-27}	1.00727	+1	$+1.60218 \times 10^{-19}$
Neutron	1.67493×10^{-27}	1.00866	0	0
Electron	0.00091×10^{-27}	0.00055	-1	-1.60218×10^{-19}

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- The number of protons in an atom is its **atomic number** and is designated by symbol Z. Each element identified by its unique atomic number, is represented by a unique chemical symbol, a one- or two-letter abbreviation.
- The sum of the protons and neutrons in an atom is called **mass number**, and is represented by symbol A.
- Atoms of the same element (same atomic number) can possess different number of neutrons (different mass numbers) and are called **isotopes**. Most elements have several isotopes, which are indicated by its chemical symbol, followed by a dash and the mass number of isotope. For example, the 3 isotopes of neon are shown below:

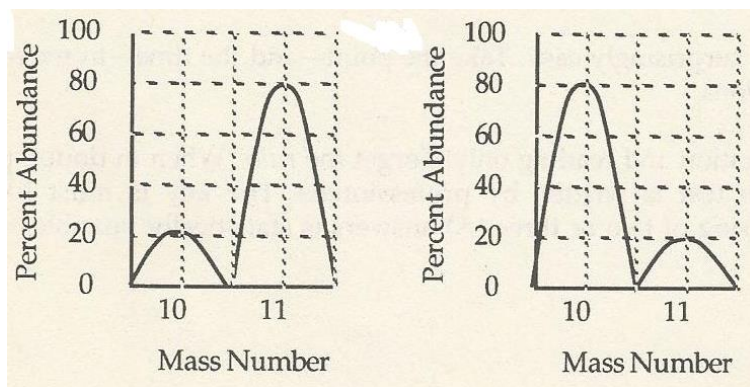
Symbol	Number of Protons	Number of Neutrons	A (Mass Number)	Natural Abundance (%)
Ne-20 or ${}^{20}_{10}\text{Ne}$	10	10	20	90.48
Ne-21 or ${}^{21}_{10}\text{Ne}$	10	11	21	0.27
Ne-22 or ${}^{22}_{10}\text{Ne}$	10	12	22	9.25

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- The relative amount of each isotope in a naturally occurring sample of a given element is represented by a percentage and is called **natural abundance**.
- The average atomic mass of an element (commonly abbreviated as atomic mass) is the weighted average of the mass of the isotopes present for this element and is closest to the most abundant isotope.

Examples:

- In light of the nuclear model of the atom, which statement below is true?
 - For a given element, the size of an isotope with more neutrons is larger than one with fewer neutrons.
 - For a given element, the size of the atom is the same for all the element's isotopes.
- Determine the number of protons and neutrons in each isotope shown below:
 - ^{222}Rn
 - ^{99}Tc
- Ca-40, K-39 and Sc-41 all have the same
 - atomic number
 - atomic mass
 - number of electrons
 - number of neutrons
- Boron has two isotopes with masses of 10 and 11 amu. Which mass diagram below represents the abundance of naturally occurring boron?



(a)

(b)

IONS & THE PERIODIC TABLE

- Main-group elements tend to form ions that have the same number of electrons as the nearest noble gas. A main-group metal tends to lose electrons, forming a *cation*, while a main-group nonmetal tends to gain electrons, forming an *anion*.

Examples:

1. Predict the charge of the ion formed by each element listed below:
 - a) Mg
 - b) N
 - c) F
 - d) Na

2. Determine the number of protons and electrons in each ion listed below:
 - a) Al^{3+}
 - b) Se^{2-}
 - c) Sr^{2+}

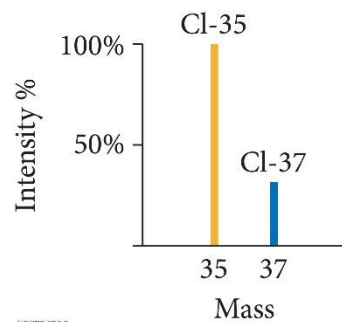
- 3.

ATOMIC MASS OF ELEMENTS

- The average atomic mass of an element (referred to as atomic mass) can be calculated from its isotopic data by using the equation below:

$$\text{Atomic mass} = \sum_n \left(\text{fraction of isotope } n \right) \times (\text{mass of isotope } n)$$

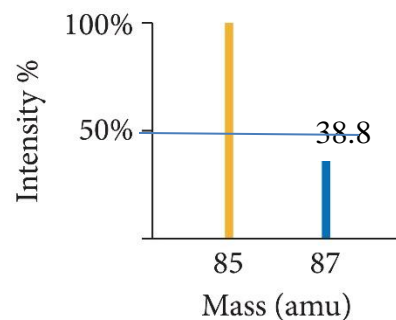
- The mass of the atom and the percent abundance of its isotopes are measured using mass spectrometry, a technique that separates particles according to their mass. The mass-spectrum of an element (shown below) indicates the abundance of each isotope relative to the most abundant isotope (usually set at 100%).



Examples:

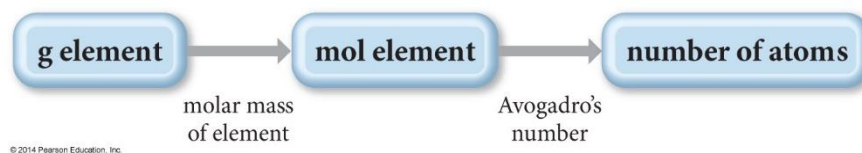
- An element has two naturally occurring isotopes. Isotope 1 has a mass of 120.9038 amu and abundance of 57.4%, and isotope 2 has a mass of 122.9042 amu. Find the atomic mass of this element and identify it.

- Use the mass spectrum of rubidium to determine the atomic mass of rubidium. (Estimate your values to 3 significant figures)



MOLE CALCULATIONS

- Mole, the counting unit in chemistry, is the amount of material containing 6.02×10^{23} particles.
- The conversions between mass, mole and particles of an element can be summarized as shown below:

**Examples:**

1. Without doing calculations, determine which sample below contains the greatest number of atoms?
 - a) 1.0 g of Fe
 - b) 1.0 g of Pb
 - c) 1.0 g of Na
2. How many helium atoms are there in a helium blimp containing 536 kg of helium?
3. Calculate the mass (in g) of one silver atom.
4. A titanium cube contains 2.86×10^{23} atoms. What is the edge length of the cube (in cm)? (The density of titanium is 4.50 g/cm^3)

Answers to In-Chapter Problems:

<i>Page</i>	<i>Example No.</i>	<i>Answer</i>
1	1	105 kg
	2	2060 g F
3	1	b
	2a	$p^+ = 86$ $n^0 = 136$
	2b	$p^+ = 43$ $n^0 = 56$
	3	d
	4	a
4	1	a) Mg^{2+} b) N^{3-} c) F^- d) Na^+
	2	a) $p^+ = 13$ $e^- = 10$ b) $p^+ = 34$ $e^- = 36$ c) $p^+ = 38$ $e^- = 36$
5	1	121.76 amu Sb
	2	85.5 amu
6	1	c
	2	8.07×10^{28} atoms
	3	1.79×10^{-22} g
	4	1.72 cm