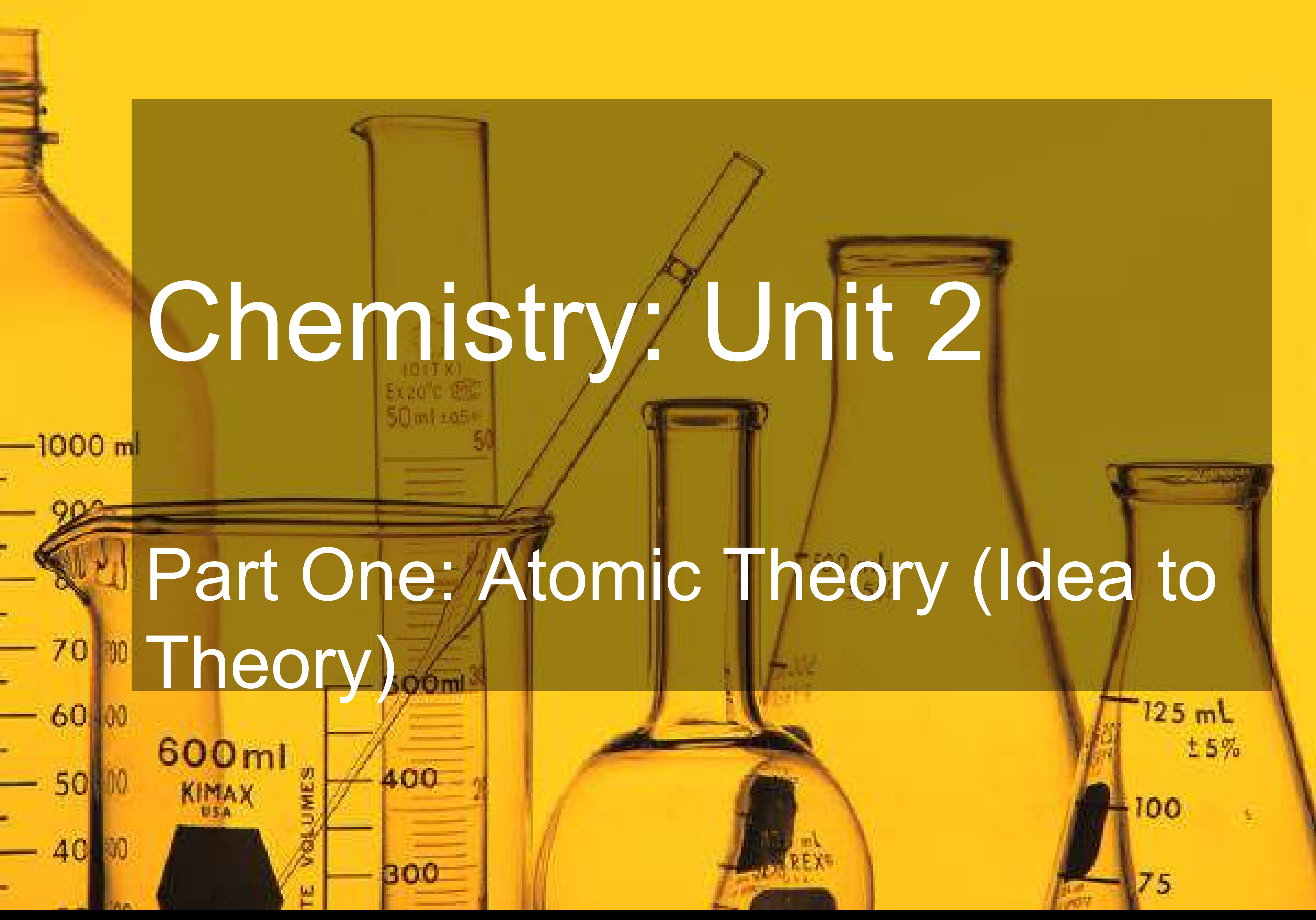


Chemistry: Unit 2

Part One: Atomic Theory (Idea to Theory)

I. Foundations of the Atomic Theory

- A. 400 B.C. –Democritus
  - 1. Defined nature’s basic particles as *atoms*
    - a. means indivisible
- B. 350 B.C. Aristotle
  - 1. said matter was made of the 4 “elements”
- C. Although these theories were wrong, they persisted for over 2000 years.

B. By the 1700s

- 1. The following ideas were accepted
  - a. elements combine to form compounds
  - b. elements could not be further broken down by ordinary means
  - c. chemical reactions are the transformations that change substances into NEW substances.

C. Basic Laws

- 1. These improved the understanding of matter and the structure of the atom.
- 2. <u>Law of Conservation of Mass</u>:
  - a. *mass is neither destroyed nor created*
- 3. <u>Law of Definite Proportions</u>:
  - a. *A chemical compound contains the same elements in exactly the same proportions by mass regardless of the sample size*

- 4. Law of Multiple Proportions:
  - a. two elements may also combine to form more than one compound
  - b. the proportions would be different for each compound

Applications of the Laws

- *Practice Problems NEEDED*
- These laws do have mathematical applications.
- These applications deal mainly with ratios and proportions

Practice Problems

- Element A – atomic mass of 2 mass units
- Element B – atomic mass of 3 mass units
- What is the expected mass of AB?

- What is the expected mass of A_2B_3?

Practice Problem

- A test tube starts out with 15.00 g of mercury (II) oxide. After heating the test tube, you find NO mercury (II) oxide left and 1.11 g of oxygen gas. What mass of liquid mercury was produced by this chemical reaction?

Practice Problem

- **If 3 g of element C combine with 8 g of element D to form compound CD, how many grams of D are needed to form compound CD_2?**

Practice Problem

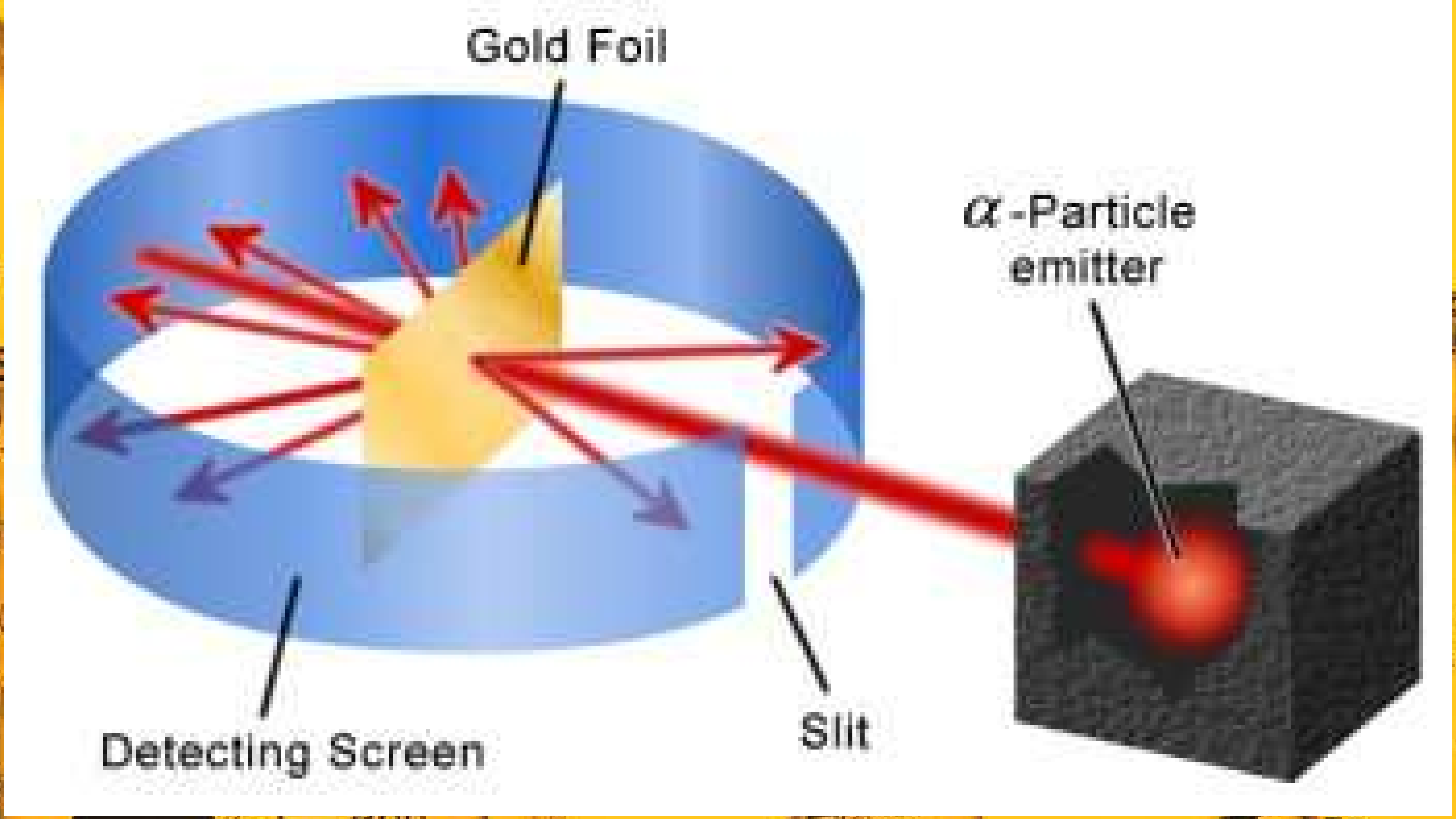
- The atomic mass of carbon is 12, and the atomic mass of oxygen is 16. To produce CO, 16 g of oxygen can be combined with 12 g of carbon. What is the ratio of oxygen to carbon when 32 g of oxygen combine with 12 g of carbon?

D. John Dalton

- 1. model was based on experimentation not pure reason
- 2. Main points
  - a. All matter is made of atoms
  - b. Atoms of an element are identical
  - c. Each element has different atoms.
  - d. Atoms of different elements combine in constant ratios to form compounds
  - e. Atoms are rearranged in reactions

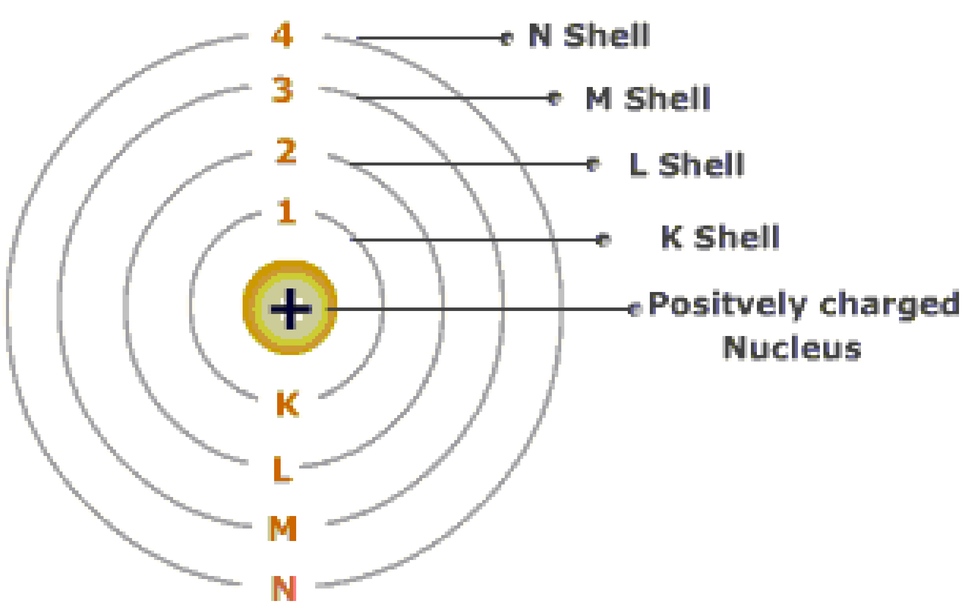
E. Rutherford

- 1. Shot alpha particles at gold foil
  - a. most passed through
    - 1. therefore, atoms are mostly empty
  - b. Some positive alpha particles bounced back
    - 1. a “nucleus” is positive & holds most of the atom’s mass


Gold Foil
α-Particle emitter
Slit
Detecting Screen

F. Bohr

- 1. Electrons orbit the nucleus in shells
- 2. Electrons can be bumped up to a higher level



G. Modern Theory

- 1. Current knowledge
  - a. Atoms are divisible into even smaller particles
- 2. Modifications
  - a. all matter is composed of atoms
  - b. atoms of any one element differ in properties from atoms of another element remain unchanged

Part II

Part II

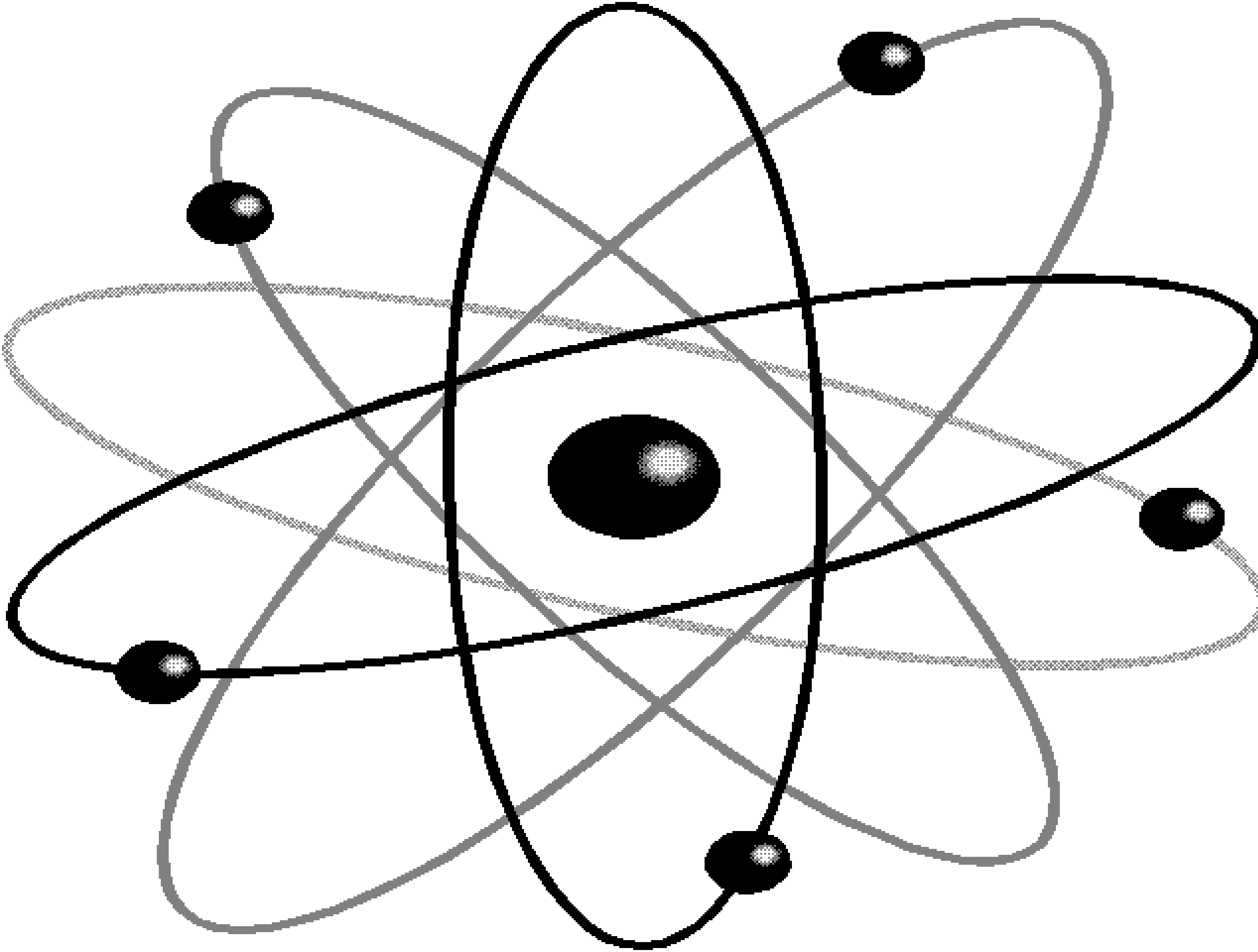
ATOMIC STRUCTURE

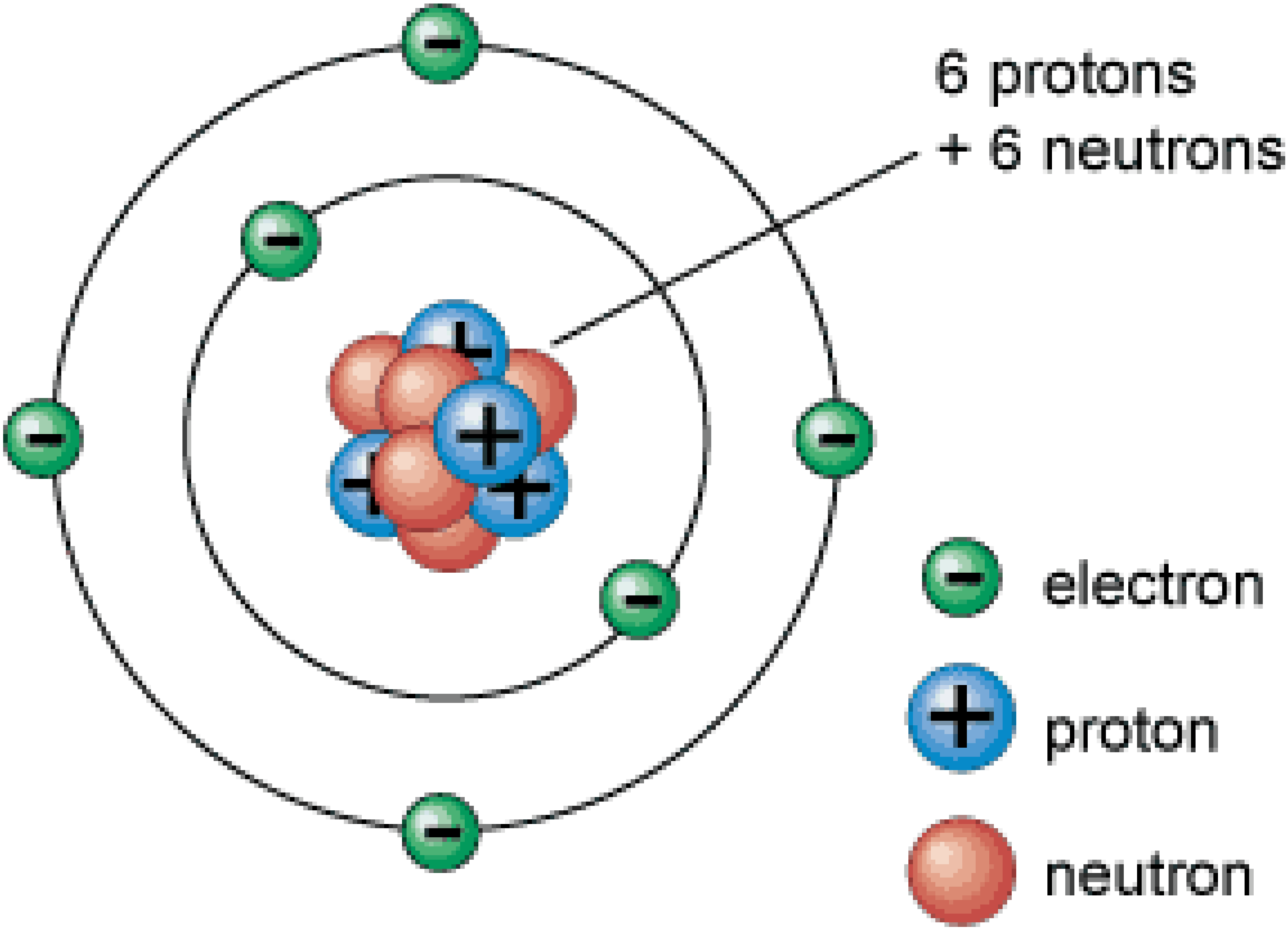
II. What are atoms made of?

- A. Nucleus
- 1. dense, center of the atom that contains protons and neutrons
- B. Neutrons
- 1. neutral charge
- 2. found in nucleus
- 3. isotopes

- C. Protons
- 1. positive charge
- 2. found in nucleus
- 3. number of protons determines atomic number of an element

- D. Electrons
- 1. negative charge
- 2. tiny particles
- 3. orbit nucleus in energy levels (electron orbitals)
- 4. ions





Carbon atom

III. The Atom and The Periodic Table

- A. Atomic number
- 1. equals the number of protons in an atom of an element
  - Different for each element
- 2. Atoms of different elements have different numbers of protons
- 3. Equals the number of electrons in a neutral atom
  - Positive Charge (protons) = Negative Charge (electrons)

B. Mass Number

- 1. Sum of the protons and the neutrons in the nucleus of an atom
  - Protons + Neutrons = Mass Number
- 2. To find the number of neutrons:
  - Mass number – atomic number = Number of neutrons

Atomic Structure on the Periodic Table

6	Atomic Number
C	Chemical Symbol
Carbon	Element Name
12.011	Atomic Mass

C. Isotopes

- 1. Atoms of the same element with different numbers of neutrons
  - Also, different mass numbers
- 2. Some isotopes are more common than others
  - Mass number on periodic table is a weighted average
  - Round the mass number from the periodic table when calculating neutron number.

D. Electrons

- 1. Electron Cloud
- a. Divided into energy levels
  - Also called, electron shells or orbitals
  - Each level holds a certain number of electrons
  - Vary in amount of energy
  - Closer to the nucleus, less energy

- b. 4 kinds
  - s orbital—simple sphere (low energy)
  - p orbital—dumbell
  - d orbital
  - f orbital--(high energy)
- c. electrons usually occupy the lowest energy levels available (stay close to nucleus)

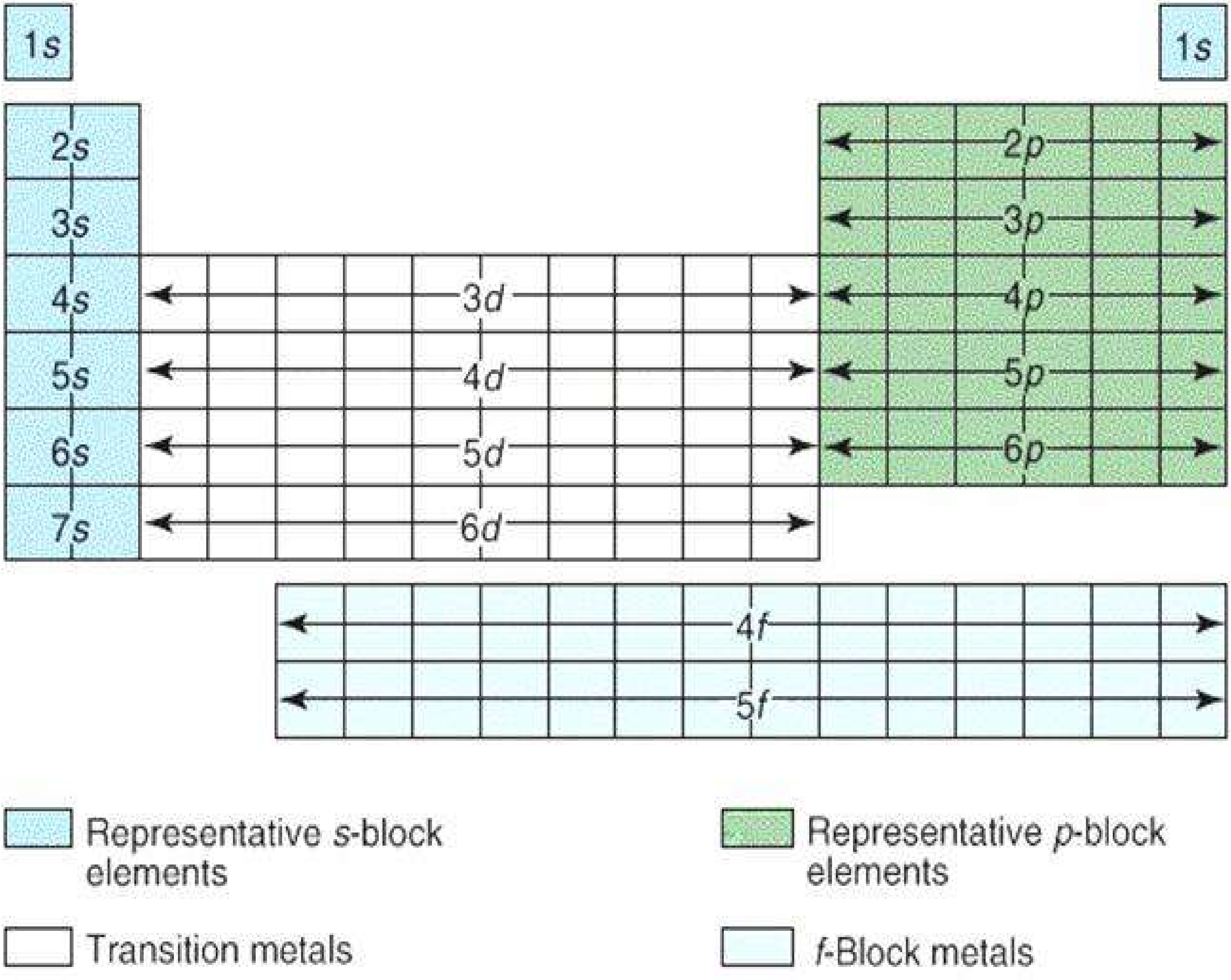
- d. Numbers of electrons in each level

Energy Level	Number of Electrons
1st	2
2nd	8
3rd	18
4th	32

2. Valence Electrons

- a. found in the outermost energy level
- b. determine how the element will react with other elements
- c. A full valence shell for all elements is 8.
  - Octet Rule
  - except Hydrogen and Helium – 2 valence electrons


germanium
www.webelements.com
1000 ml
25 mL
± 5%



1s
1s
2s
3s
4s
5s
6s
7s
3d
4d
5d
6d
2p
3p
4p
5p
6p
4f
5f
Representative s-block elements
Representative p-block elements
Transition metals
f-Block metals

E. Relative Atomic Mass

- 1. Scientists use one atom, Carbon-12 as the standard for atomic mass
- 2. 1 atomic mass unit (amu) = 1/12 the mass of the a carbon-12 atom
- 3. Mass number and relative atomic mass are very close, but not identical

F. Average Atomic Mass

- 1. The weighted average of the atomic masses of NATURALLY occurring isotopes of an element.
  - Mass given on the period table
- 2.

$$\text{Avg. Atomic Mass} = \frac{(\text{mass})(\%) + (\text{mass})(\%)}{100}$$

3. Calculating Average Atomic Mass

- Naturally occurring copper consists of 69.17% Copper-63 (Atomic mass = 62.93 amu) and 30.83% Copper-65 (Atomic mass = 64.93 amu). Calculate the average atomic mass.

Average Atomic Mass Practice

- <u>EX</u>: Calculate the avg. atomic mass of oxygen if its abundance in nature is 99.76% ^{16}O, 0.04% ^{17}O, and 0.20% ^{18}O.

$$\text{Avg. Atomic Mass} = \frac{(16)(99.76) + (17)(0.04) + (18)(0.20)}{100} = 16.00 \text{ amu}$$

Average Atomic Mass Practice

- <u>EX</u>: Find chlorine's average atomic mass if approximately 8 of every 10 atoms are chlorine-35 and 2 are chlorine-37.

$$\text{Avg. Atomic Mass} = \frac{(35)(8)+(37)(2)}{10} = 35.40 \text{ amu}$$

G. The MOLE and Atomic Mass

- 1. Mole = SI unit for the amount of a substance
  - Contains as many particles as there are atoms in exactly 12 g of carbon-12
  - This is a counting unit (like a dozen)
  - 6.022 x 10^{23} carbon-12 atoms = 12 grams of carbon-12

2. Avogadro’s Number

- a. the number of particles in exactly one mole of a pure substance
- b. 6.022×10^{23} particles (Learn it!)

3. Molar Mass

- A. The mass of one mole of a pure substance
- B. Unit = g/mol
- C. molar mass is numerically equal to the atomic mass of the element in atomic mass units (on periodic table)

D. Gram-Mole Conversions

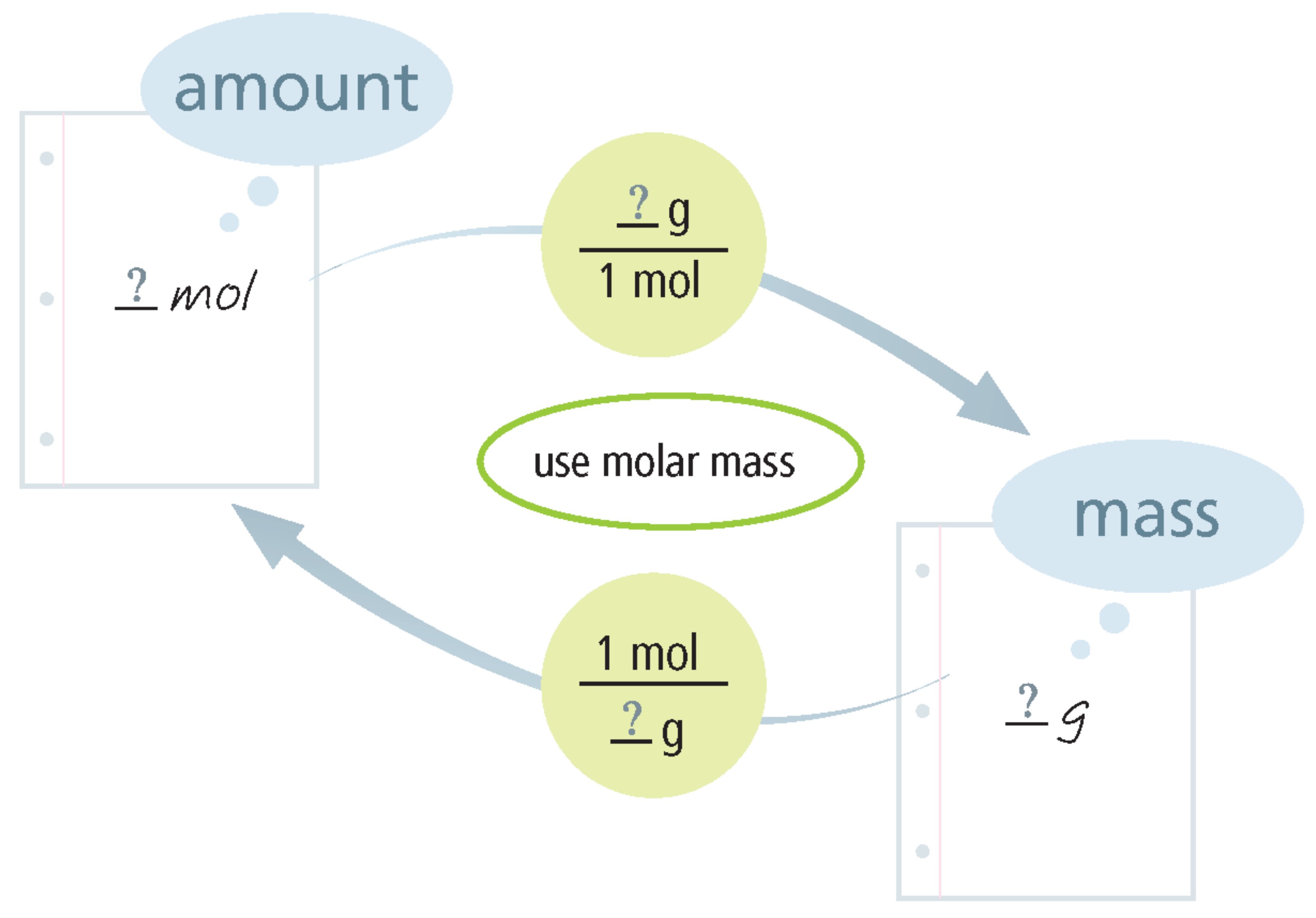
- **The conversion factor for gram-mole conversion is molar mass.**

$$\frac{g}{mol} \quad \text{OR} \quad \frac{mol}{g}$$

Practice Problems page 85

1. **What is the mass in grams of 2.25 mol of the element iron?** 126 g Fe
2. **What is the mass in grams of 0.357 mol of the element potassium?** 14.7 g K
3. **What is the mass in grams of 0.0135 mol of the element sodium?** 0.310 g Na
4. **What is the mass in grams of 16.3 mol of the element nickel?** 957 g Ni

Conversions Image



p. 84

Gram-Mole Conversions

- **The conversion factor for gram-mole conversion is molar mass.**

$$\frac{g}{mol} \quad \text{OR} \quad \frac{mol}{g}$$

- **A Chemist produced 11.9 g of Al. How many moles of Al were produced?**
  - **0.411 moles Al**

Practice Problems page 85

1. **How many moles of calcium are in 5.00 g of calcium?**
   0.125 mol Ca
2. **How many moles of gold are in 3.60 x 10^{-5} g of gold?**
   1.83 x 10^{-7} mol Au
3. **How many moles of zinc are in 0.535 g of zinc?**
   8.18 x 10^{-3} mol Zn

Conversions with Avogadro's Number

- The conversion factor for particle-mole conversion is Avogadro's number:

$$\frac{6.022\times10^{23}\ \text{atoms}}{1\ \text{mol}} \quad \text{OR} \quad \frac{1\ \text{mol}}{6.022\times10^{23}\ \text{atoms}}$$

- How many moles of silver are in 3.01×10^{23} atoms of silver
  - 0.500 moles Ag

Practice Problems page 86

1. **How many moles of lead are 1.50 x 10^{12} atoms of lead?** 2.49×10^{-12} mol Pb
2. **How many moles of tin are in 2500 atoms of tin?** 4.2×10^{-21} mol Sn
3. **How many atoms of aluminum are in 2.75 mol of aluminum?**

1.66×10^{24} atoms Al

Conversions with Avogadro's Number

- **The conversion factor for particle-mole conversion is Avogadro's number.**

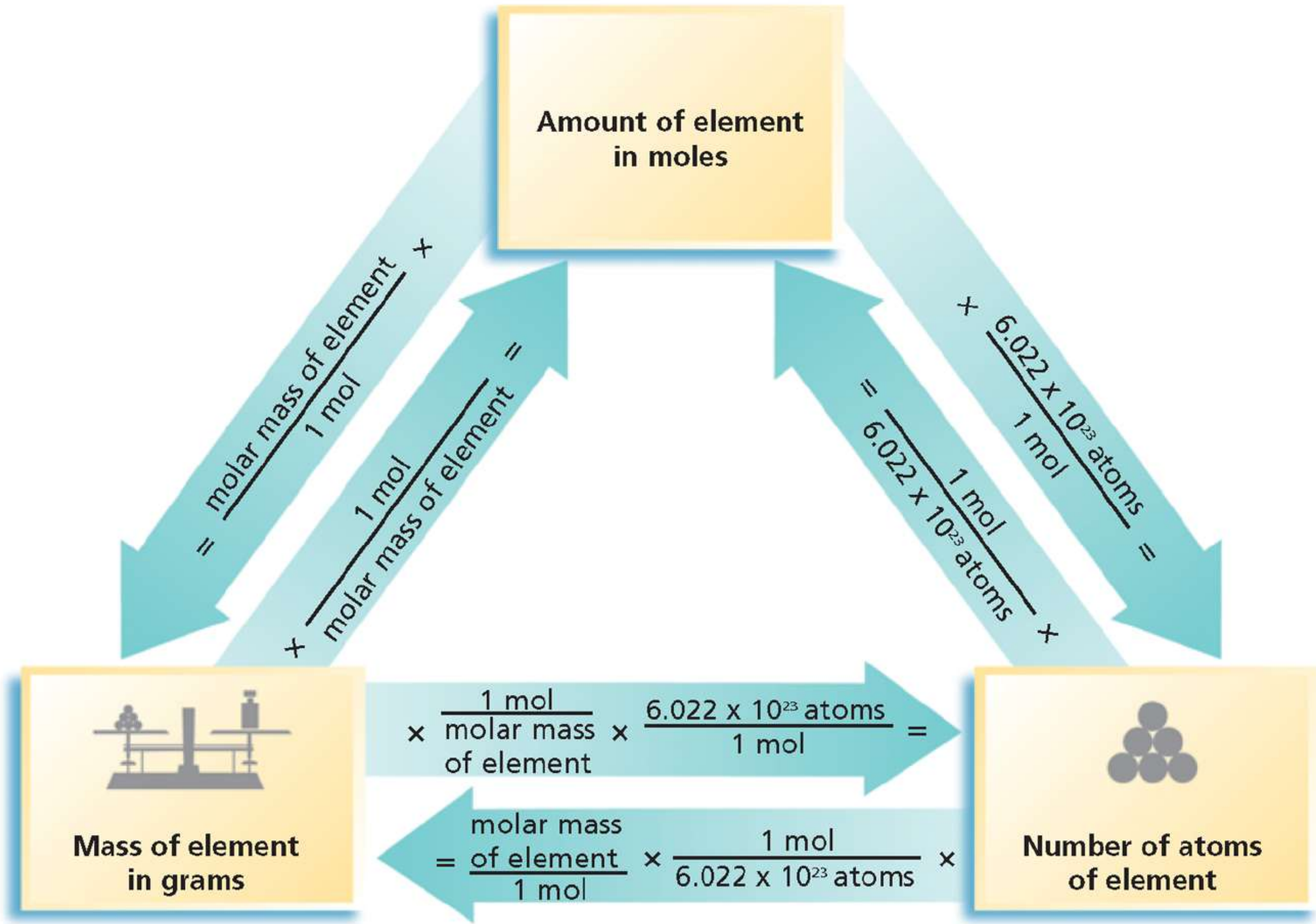
$$\frac{6.022\text{x}10^{23}\text{ atoms}}{1\text{ mol}} \quad \text{OR} \quad \frac{1\text{ mol}}{6.022\text{x}10^{23}\text{atoms}}$$

- **What is the mass, in grams, of $1.20\text{x}10^{18}$ atoms of Cu?**
  - **$1.27 \text{ x } 10^{-4}$ g Cu**

Practice Problems page 87

1. **What is the mass in grams of 7.5×10^{15} atoms of nickel?**
   7.3×10^{-7} g Ni
2. **How many atoms of sulfur are in 4.00 g of sulfur?**
   7.51×10^{22} atoms S
3. **What mass of gold contains the same number of atoms as 9.0 g of aluminum?**
   66 g Au

Conversions Image



III. Application to compounds

- A. Molar Mass of Compounds
  - 1. Also known as molecular mass or formula mass
  - 2. To calculate
    - Simple add all the atomic masses for each element in the chemical formula

3. Practice Problems

- 1. NaBr (sodium bromide)
- 2. HNO_3 (nitric acid)
- 3. H_2O (Water)
- 4. NaOH (sodium hydroxide)
- 5. $CaCO_3$ (calcium carbonate)

B. Percent Composition

- 1. The mass percent of an element in a compound is calculated from the mass of the element present in one mole of the compound divided by the mass of one mole of the compound and converted to a percent.

2. Formula

$$\text{Mass \% of element} = \frac{\text{atomic mass of element in the compound}}{\text{formula weight of compound containing that element}}$$

3. Practice Problems

- A. What is the percent composition of Hg in HgO?
- B. What is the percent composition of each element in Na_3PO_4?

- C. What is the percent composition of S in SO_2?

- D. How much mercury would be obtained by decomposing 7.65 g of HgO?