Preview

Lesson Starter

- Young people should not smoke.
- Smoking at an early age may make it more difficult to quit smoking later.
- Which of the above statements is an opinion and which is a theory?
- Which is similar to Aristotle's statements?

End Of Slide

Objectives

- **Explain** the *law of conservation of mass,* the *law of definite proportions,* and the *law of multiple proportions.*
- **Summarize** the five essential points of Dalton's atomic theory.
- **Explain** the relationship between Dalton's atomic theory and the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.

End Of Slide

Foundations of Atomic Theory

- The transformation of a substance or substances into one or more new substances is known as a *chemical reaction.*

- **Law of conservation of mass:** mass is neither created nor destroyed during ordinary chemical reactions or physical changes

End Of Slide

Chemical Reaction

Click below to watch the Visual Concept.

Law of Conservation of Mass

Click below to watch the Visual Concept.

Visual Concept

Foundations of Atomic Theory, *continued*

- **Law of definite proportions:** a chemical compound contains the same elements in exactly the same proportions by mass regardless of the size of the sample or source of the compound
- **Law of multiple proportions:** if two or more different compounds are composed of the same two elements, then the ratio of the masses of the second element combined with a certain mass of the first element is always a ratio of small whole numbers

End Of Slide

Law of Definite Proportions

Click below to watch the Visual Concept.

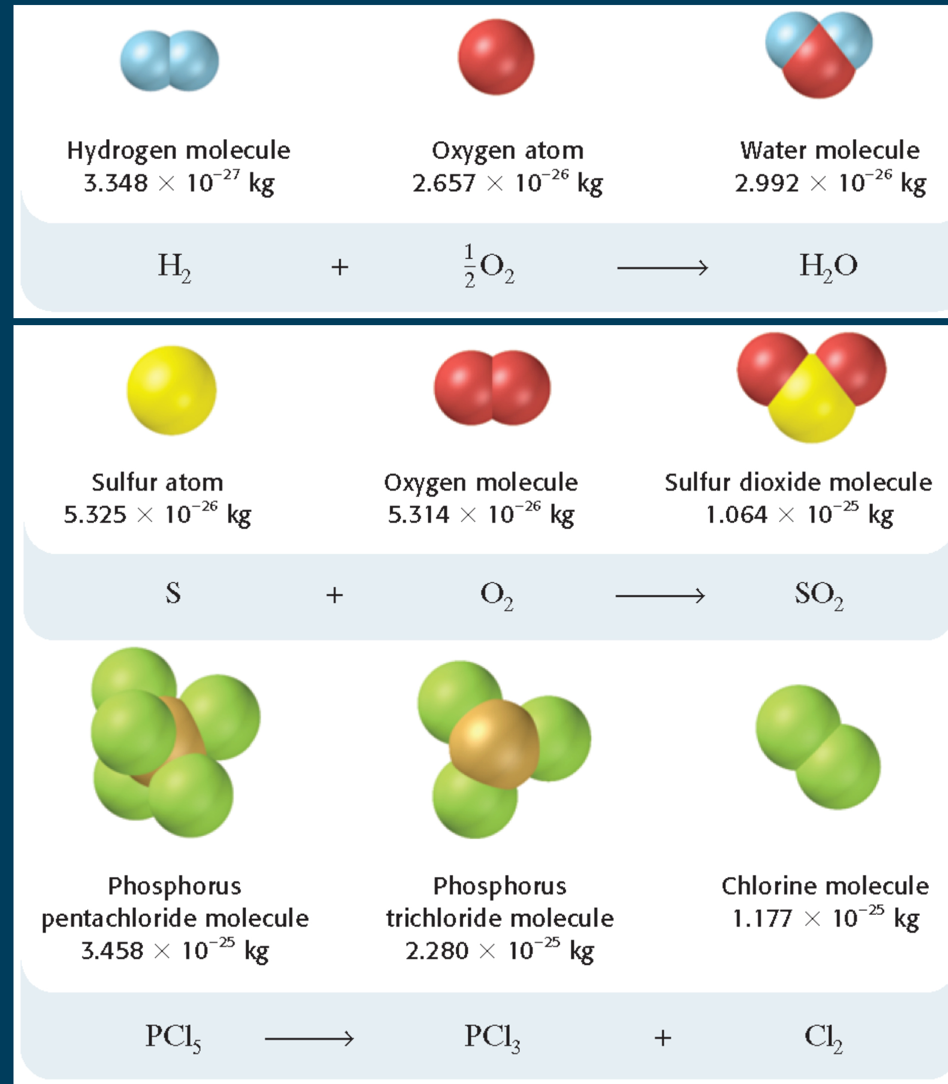
Visual Concept

Law of Multiple Proportions

Click below to watch the Visual Concept.

Visual Concept

Law of Conservation of Mass



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Law of Multiple Proportions

Name of compound	Description	As shown in figures	Formula	Mass O (*g*)	Mass N (*g*)	$\frac{\text{Mass O (g)}}{\text{Mass N (g)}}$
Nitrogen monoxide	colorless gas that reacts readily with oxygen		NO	16.00	14.01	$\frac{16.00 \text{ g O}}{14.01 \text{ g N}} = \frac{1.14 \text{ g O}}{1 \text{ g N}}$
Nitrogen dioxide	poisonous brown gas in smog		NO_2	32.00	14.01	$\frac{32.00 \text{ g O}}{14.01 \text{ g N}} = \frac{2.28 \text{ g O}}{1 \text{ g N}}$

Dalton's Atomic Theory

- All matter is composed of extremely small particles called atoms.
- Atoms of a given element are identical in size, mass, and other properties; atoms of different elements differ in size, mass, and other properties.
- Atoms cannot be subdivided, created, or destroyed.

End of Slide

Dalton's Atomic Theory, *continued*

- Atoms of different elements combine in simple whole-number ratios to form chemical compounds.

- In chemical reactions, atoms are combined, separated, or rearranged.

End Of Slide

Modern Atomic Theory

- Not all aspects of Dalton's atomic theory have proven to be correct. We now know that:
  - Atoms are divisible into even smaller particles.
  - A given element can have atoms with different masses.
- Some important concepts remain unchanged.
  - All matter is composed of atoms.
  - Atoms of any one element differ in properties from atoms of another element.

Preview

Lesson Starter

- Even though the two shapes look different, the characteristics of the various parts that compose them are the same.
- The same is true with the atom.
- Though atoms of different elements display different properties, isolated subatomic particles have the same properties.

Objectives

- **Summarize** the observed properties of cathode rays that led to the discovery of the electron.
- **Summarize** the experiment carried out by Rutherford and his co-workers that led to the discovery of the nucleus.
- **List** the properties of protons, neutrons, and electrons.
- **Define** *atom.*

The Structure of the Atom

- An **atom** is the smallest particle of an element that retains the chemical properties of that element.
- The *nucleus* is a very small region located at the center of an atom.
- The nucleus is made up of at least one positively charged particle called a *proton* and usually one or more neutral particles called *neutrons.*

End Of Slide

The Structure of the Atom, *continued*

- Surrounding the nucleus is a region occupied by negatively charged particles called *electrons.*

- Protons, neutrons, and electrons are often referred to as *subatomic particles.*

Atom

Click below to watch the Visual Concept.

Parts of the Atom

Click below to watch the Visual Concept.

Visual Concept

Properties of Subatomic Particles

Particle	Symbols	Relative electric charge	Mass number	Relative mass (amu*)	Actual mass (kg)
Electron	e^-, ${}^{0}_{-1}e$	–1	0	0.000 5486	9.109×10^{-31}
Proton	p^+, ${}^{1}_{1}\mathrm{H}$	+1	1	1.007 276	1.673×10^{-27}
Neutron	n°, ${}^{1}_{0}n$	0	1	1.008 665	1.675×10^{-27}

*1 amu (atomic mass unit) = $1.660\,540 \times 10^{-27}$ kg

Discovery of the Electron

Cathode Rays and Electrons

- Experiments in the late 1800s showed that cathode rays were composed of negatively charged particles.
- These particles were named *electrons.*

End of Slide

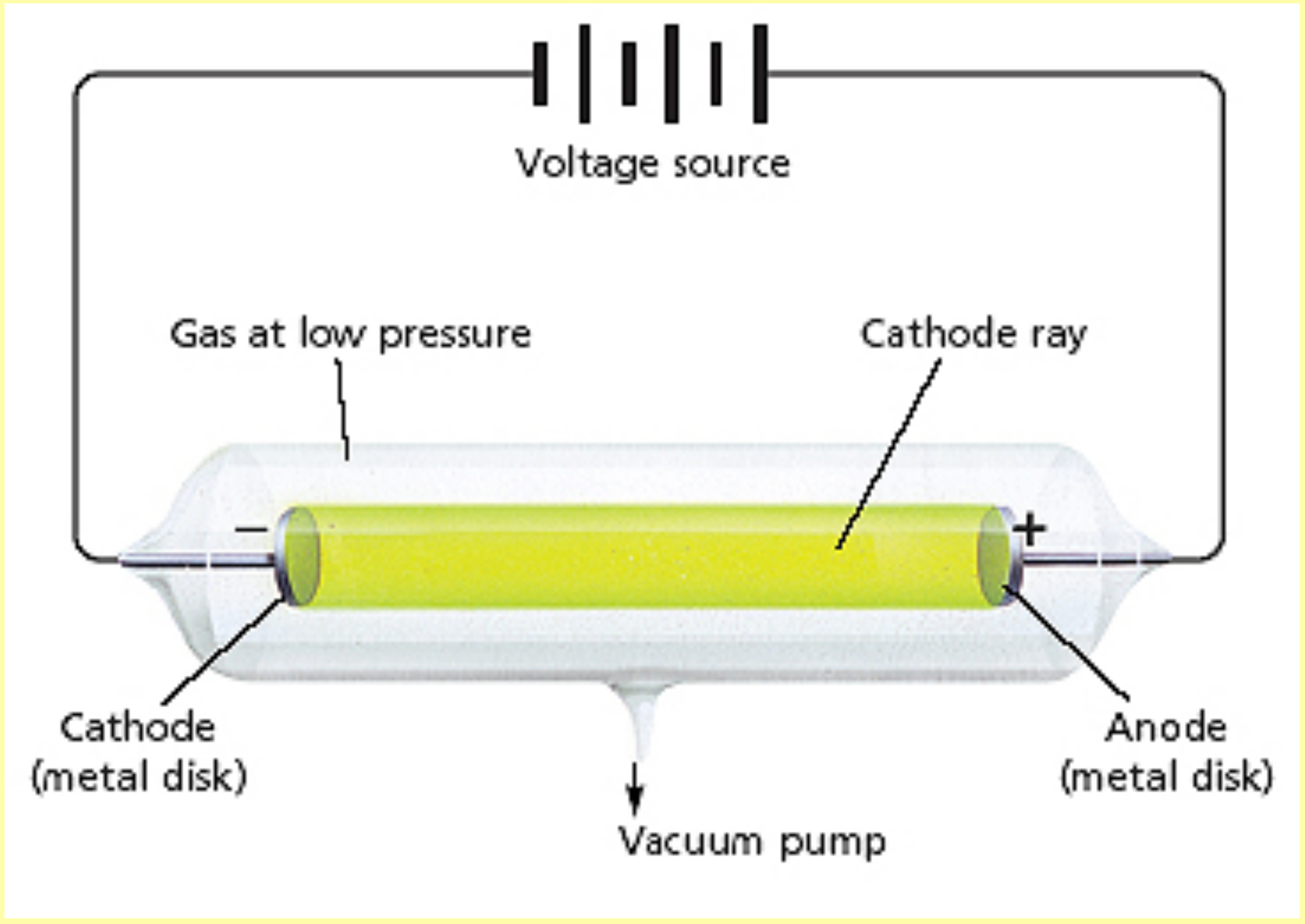
Discovery of the Electron, *continued*

Charge and Mass of the Electron

- Joseph John Thomson's cathode-ray tube experiments measured the charge-to-mass ratio of an electron.
- Robert A. Millikan's oil drop experiment measured the charge of an electron.
- With this information, scientists were able to determine the mass of an electron.

End Of Slide

Discovery of the Electron, *continued*



Thomson’s Cathode-Ray Tube Experiment

Click below to watch the Visual Concept.

Visual Concept

Millikan's Oil Drop Experiment

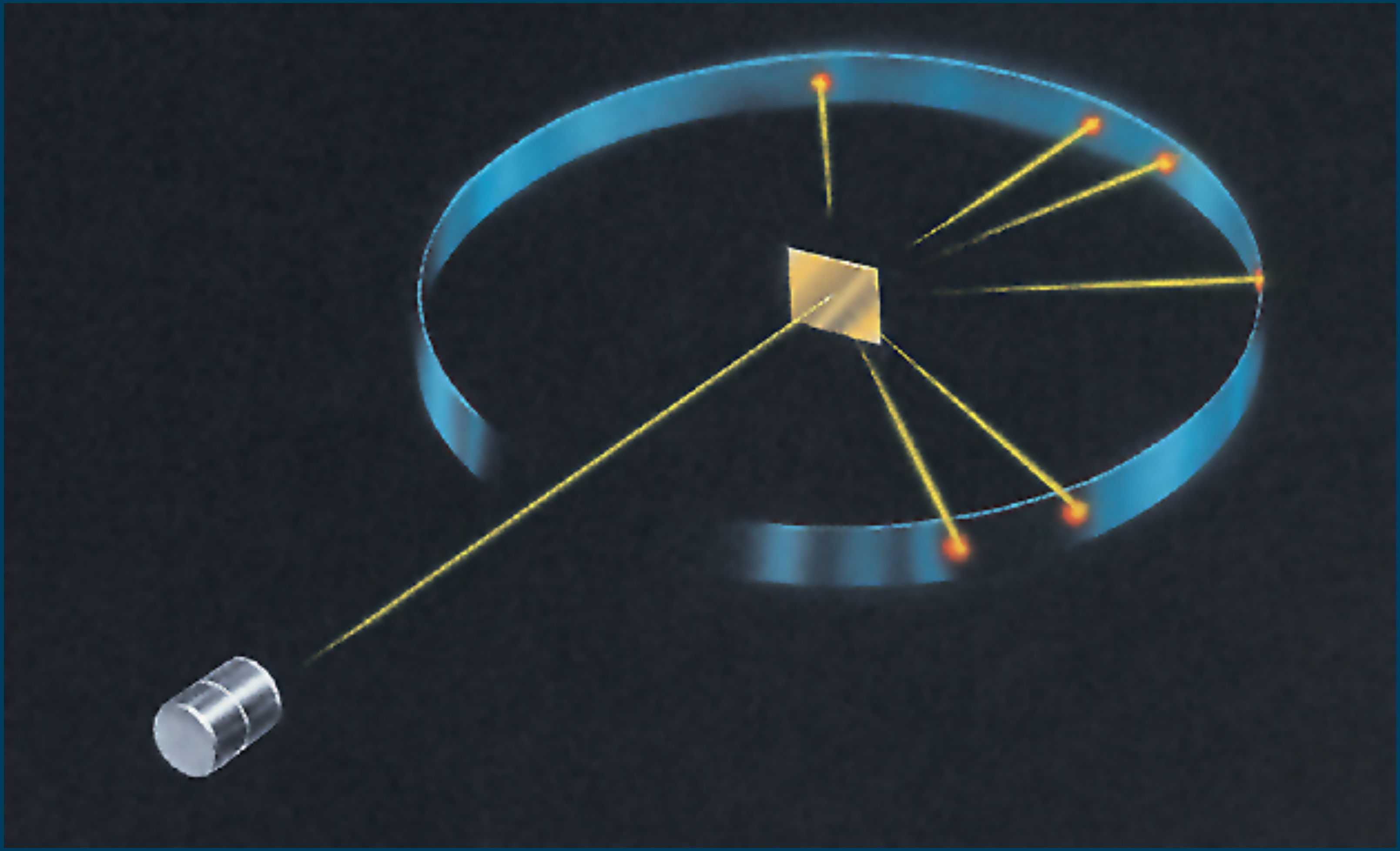
Click below to watch the Visual Concept.

Visual Concept

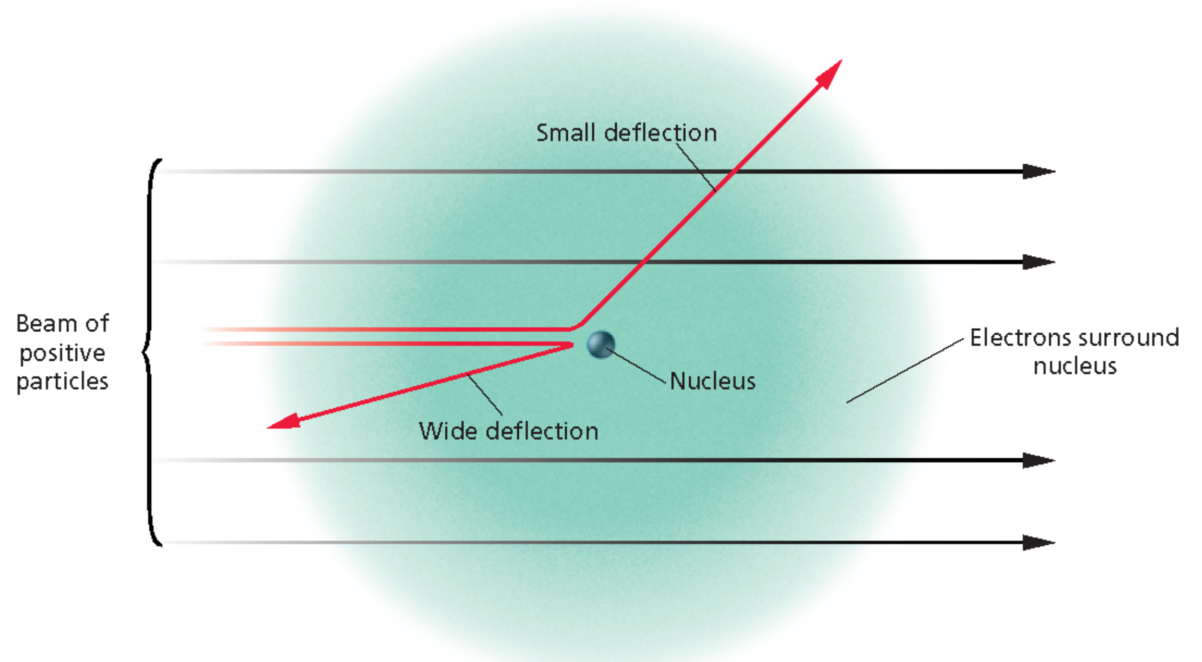
Discovery of the Atomic Nucleus

- More detail of the atom's structure was provided in 1911 by Ernest Rutherford and his associates Hans Geiger and Ernest Marsden.
- The results of their gold foil experiment led to the discovery of a very densely packed bundle of matter with a positive electric charge.
- Rutherford called this positive bundle of matter the *nucleus.*

Gold Foil Experiment



Gold Foil Experiment on the Atomic Level



Rutherford reasoned that each atom in the gold foil contained a small, dense, positively charged nucleus surrounded by electrons. A small number of the alpha particles directed toward the foil were deflected by the tiny nucleus (red arrows). Most of the particles passed through undisturbed (black arrows).

Rutherford’s Gold Foil Experiment

Click below to watch the Visual Concept.

Visual Concept

Composition of the Atomic Nucleus

- Except for the nucleus of the simplest type of hydrogen atom, all atomic nuclei are made of *protons* and *neutrons.*
- A proton has a positive charge equal in magnitude to the negative charge of an electron.
- Atoms are electrically neutral because they contain equal numbers of protons and electrons.
- A neutron is electrically neutral.

Composition of the Atomic Nucleus, *continued*

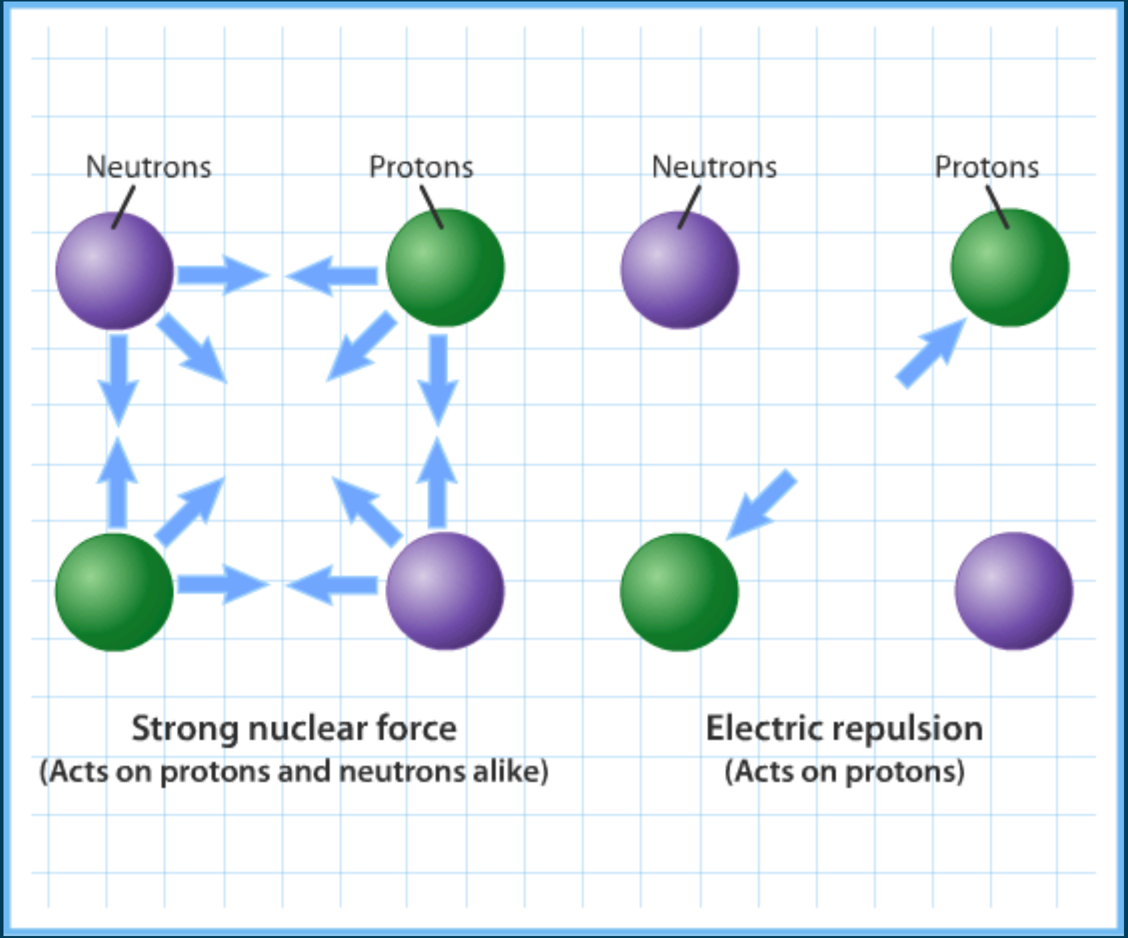
- The nuclei of atoms of different elements differ in their number of protons and therefore in the amount of positive charge they possess.
- Thus, the number of protons determines that atom's identity.

Composition of the Atomic Nucleus, *continued*

Forces in the Nucleus

- When two protons are extremely close to each other, there is a strong attraction between them.
- A similar attraction exists when neutrons are very close to each other or when protons and neutrons are very close together.
- The short-range proton-neutron, proton-proton, and neutron-neutron forces that hold the nuclear particles together are referred to as **nuclear forces.**

Nuclear Forces



The Sizes of Atoms

- The radius of an atom is the distance from the center of the nucleus to the outer portion of its electron cloud.
- Because atomic radii are so small, they are expressed using a unit that is more convenient for the sizes of atoms.
- This unit is the *picometer, pm.*

Preview

Lesson Starter

- Imagine that your semester grade depends 60% on exam scores and 40% on laboratory explorations.
- Your exam scores would count more heavily toward your final grade.
- In this section, you will learn that the atomic mass of an element is a weighted average of the masses of the naturally occurring isotopes of that element.

End Of Slide

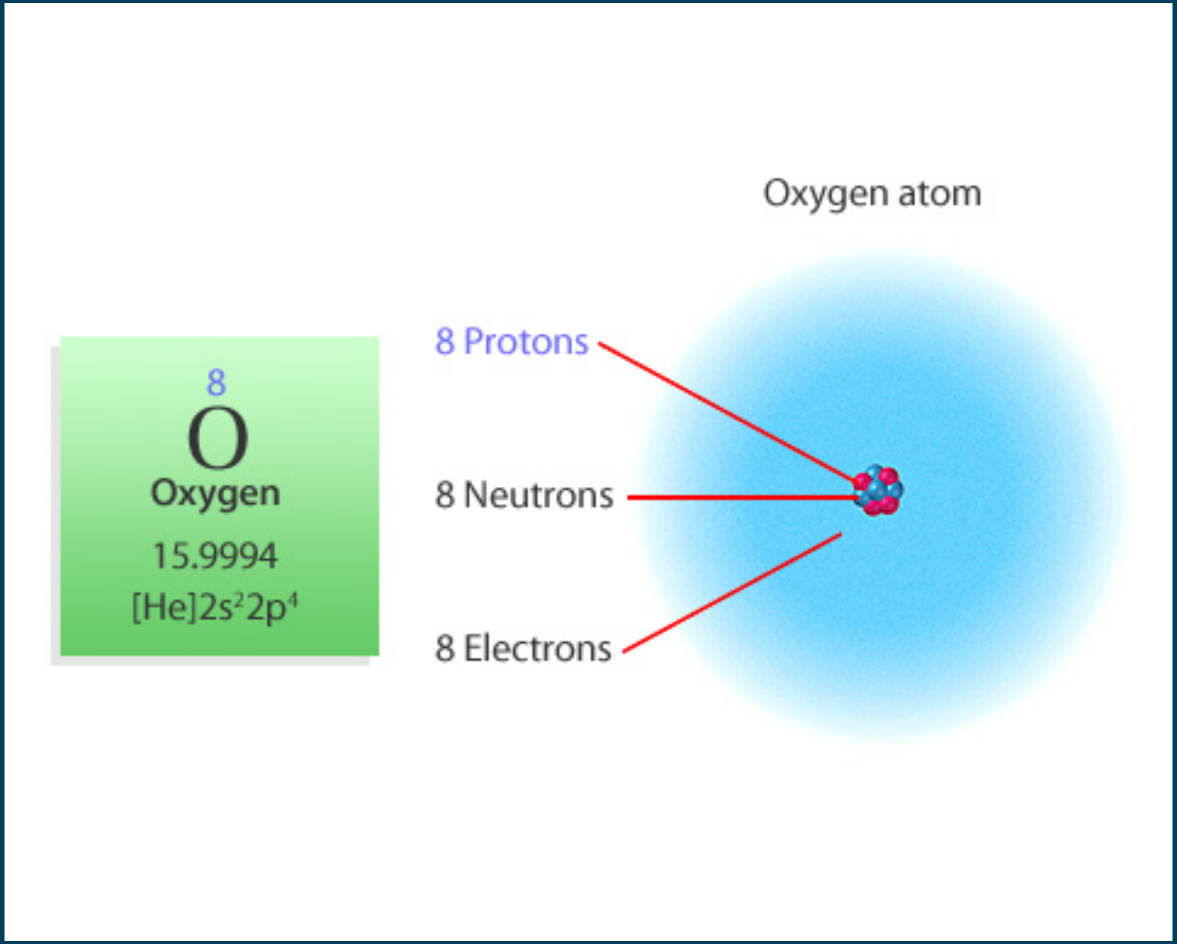
Objectives

- **Explain** what isotopes are.
- **Define** *atomic number* and *mass number,* and **describe** how they apply to isotopes.
- Given the identity of a nuclide, **determine** its number of protons, neutrons, and electrons.
- **Define** *mole, Avogadro's number,* and *molar mass,* and state how all three are related.
- **Solve** problems involving mass in grams, amount in moles, and number of atoms of an element.

Atomic Number

- Atoms of different elements have different numbers of protons.
- Atoms of the same element all have the same number of protons.
- The **atomic number** (Z) of an element is the number of protons of each atom of that element.

Atomic Number



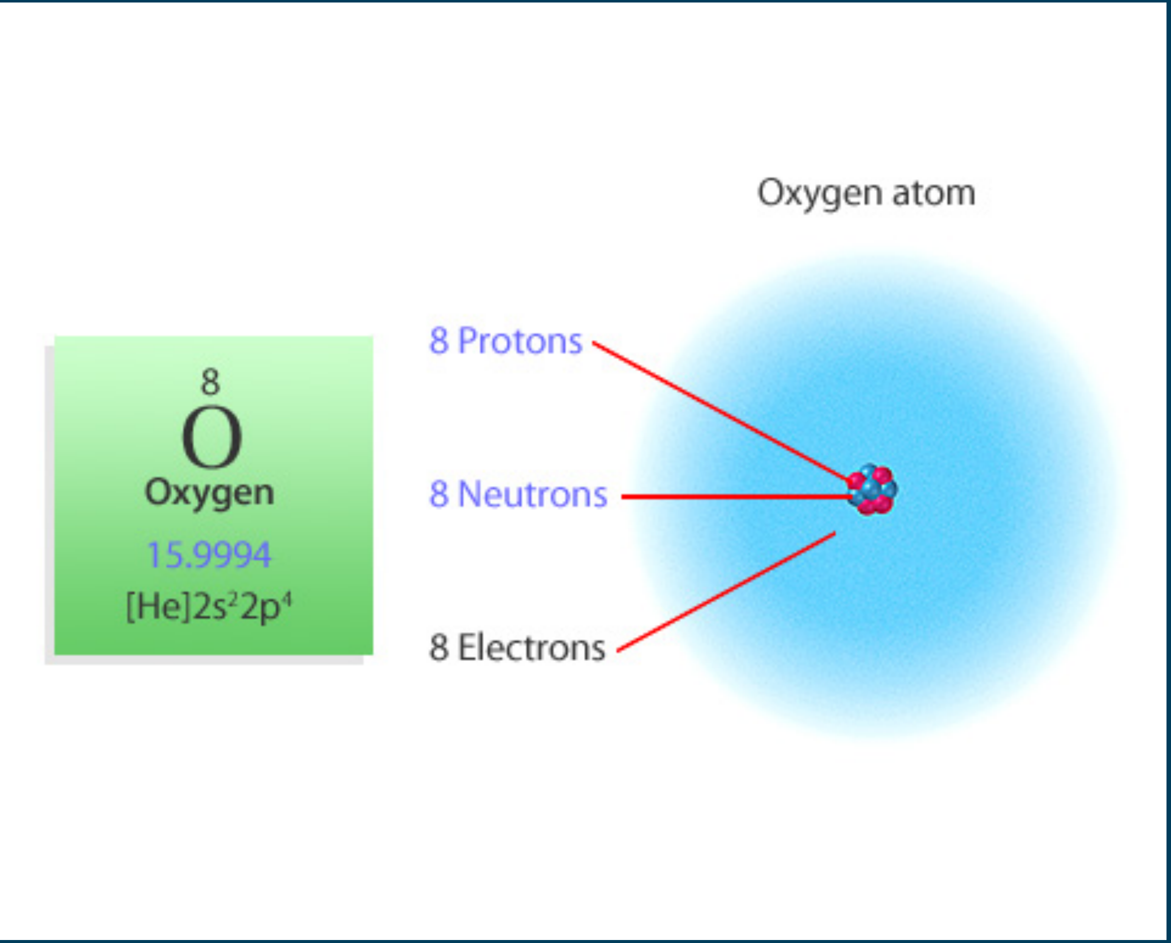
Isotopes

- **Isotopes** are atoms of the same element that have different masses.
- The isotopes of a particular element all have the same number of protons and electrons but different numbers of neutrons.
- Most of the elements consist of mixtures of isotopes.

Mass Number

- The **mass number** is the total number of protons and neutrons that make up the nucleus of an isotope.

Mass Number


Oxygen atom
8
O
Oxygen
15.9994
[He]2s²2p⁴
8 Protons
8 Neutrons
8 Electrons

End Of Slide

Designating Isotopes

- Hyphen notation: The mass number is written with a hyphen after the name of the element.
  - uranium-235
- Nuclear symbol: The superscript indicates the mass number and the subscript indicates the atomic number.

$$^{235}_{92}\mathrm{U}$$

Designating Isotopes, *continued*

- The number of neutrons is found by subtracting the atomic number from the mass number.

mass number − atomic number = number of neutrons

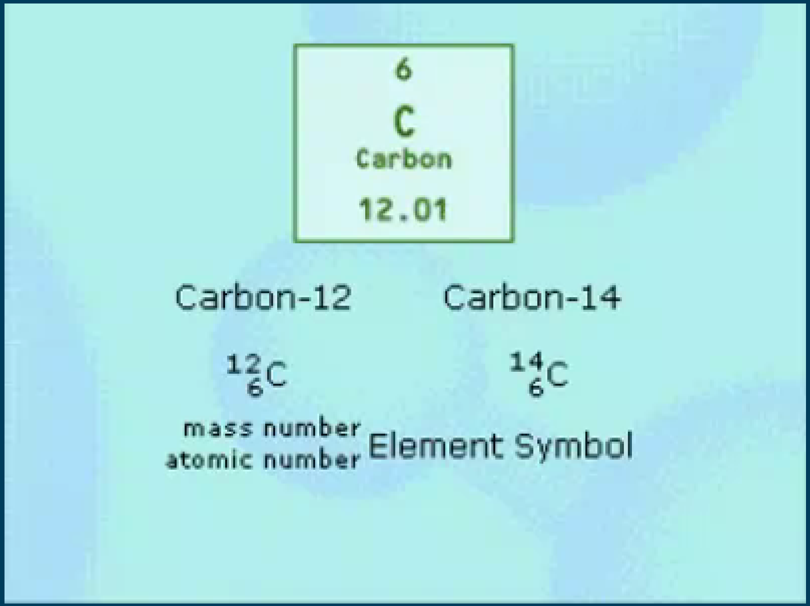
235 (protons + neutrons) − 92 protons = 143 neutrons

- **Nuclide** is a general term for a specific isotope of an element.

End Of Slide

Isotopes and Nuclides

Click below to watch the Visual Concept.



Designating Isotopes, *continued*

Sample Problem A

How many protons, electrons, and neutrons are there in an atom of chlorine-37?

End of Slide

Designating Isotopes, *continued*

Sample Problem A Solution

Given: name and mass number of chlorine-37

Unknown: numbers of protons, electrons, and neutrons

Solution:

atomic number = number of protons = number of electrons

mass number = number of neutrons + number of protons

End Of Slide

Designating Isotopes, *continued*

Sample Problem A Solution, *continued*

mass number of chlorine-37 − atomic number of chlorine = number of neutrons in chlorine-37

mass number − atomic number = 37 (protons plus neutrons) − 17 protons = 20 neutrons

An atom of chlorine-37 is made up of 17 electrons, 17 protons, and 20 neutrons.

End Of Slide

Relative Atomic Masses

- The standard used by scientists to compare units of atomic mass is the carbon-12 atom, which has been arbitrarily assigned a mass of exactly 12 atomic mass units, or 12 amu.
- One **atomic mass unit,** or 1 **amu,** is exactly 1/12 the mass of a carbon-12 atom.
- The atomic mass of any atom is determined by comparing it with the mass of the carbon-12 atom.

Average Atomic Masses of Elements

- **Average atomic mass** is the weighted average of the atomic masses of the naturally occurring isotopes of an element.

Calculating Average Atomic Mass

- The average atomic mass of an element depends on both the mass and the relative abundance of each of the element's isotopes.

End Of Slide

Average Atomic Masses of Elements, *continued*

Calculating Average Atomic Mass, *continued*

- Copper consists of 69.15% copper-63, which has an atomic mass of 62.929 601 amu, and 30.85% copper-65, which has an atomic mass of 64.927 794 amu.

- The average atomic mass of copper can be calculated by multiplying the atomic mass of each isotope by its relative abundance (expressed in decimal form) and adding the results.

End of Slide

Average Atomic Masses of Elements, *continued*

Calculating Average Atomic Mass, *continued*

- (0.6915 × 62.929 601 amu) + (0.3085 × 64.927 794 amu) = 63.55 amu
- The calculated average atomic mass of naturally occurring copper is 63.55 amu.

Average Atomic Mass

Click below to watch the Visual Concept.

Visual Concept

Relating Mass to Numbers of Atoms

The Mole

- The mole is the SI unit for amount of substance.
- A **mole** (abbreviated mol) is the amount of a substance that contains as many particles as there are atoms in exactly 12 g of carbon-12.

Avogadro's Number

- **Avogadro's number**—6.022 1415 × 10^{23}—is the number of particles in exactly one mole of a pure substance.

End Of Slide

Relating Mass to Numbers of Atoms, *continued*

Molar Mass

- The mass of one mole of a pure substance is called the molar mass of that substance.

- Molar mass is usually written in units of g/mol.

- The molar mass of an element is numerically equal to the atomic mass of the element in atomic mass units.

End Of Slide

Relating Mass to Numbers of Atoms, *continued*

Gram/Mole Conversions

- Chemists use molar mass as a conversion factor in chemical calculations.

- For example, the molar mass of helium is 4.00 g He/mol He.

- To find how many grams of helium there are in two moles of helium, multiply by the molar mass.

$$2.00 \text{ mol He} \times \frac{4.00 \text{ g He}}{1 \text{ mol He}} = 8.00 \text{ g He}$$

Relating Mass to Numbers of Atoms, *continued*

Conversions with Avogadro's Number

- Avogadro's number can be used to find the number of atoms of an element from the amount in moles or to find the amount of an element in moles from the number of atoms.
- In these calculations, Avogadro's number is expressed in units of atoms per mole.

End Of Slide

The Mole

Click below to watch the Visual Concept.

Visual Concept

Avogadro's Number

Click below to watch the Visual Concept.

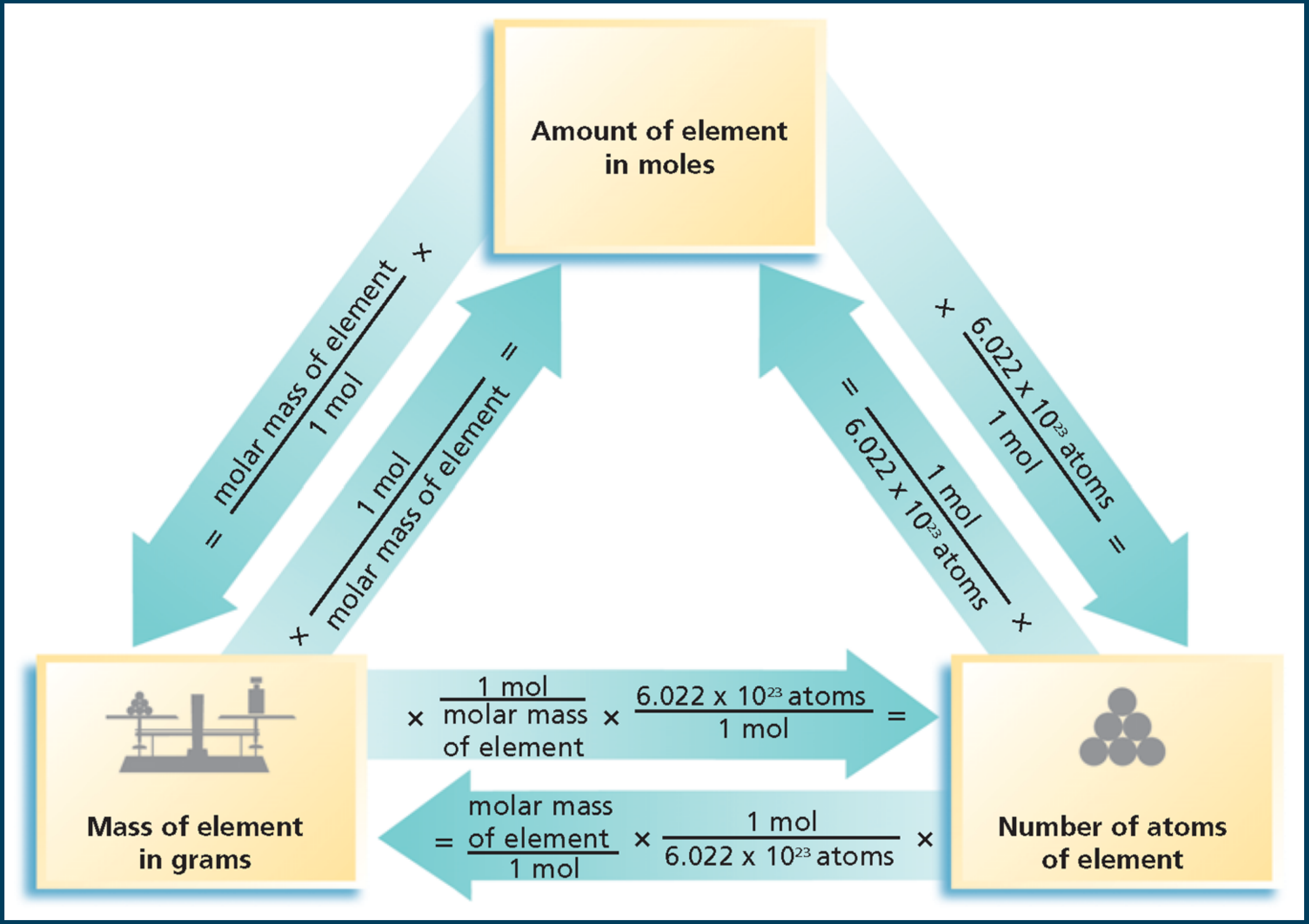
Visual Concept

Molar Mass

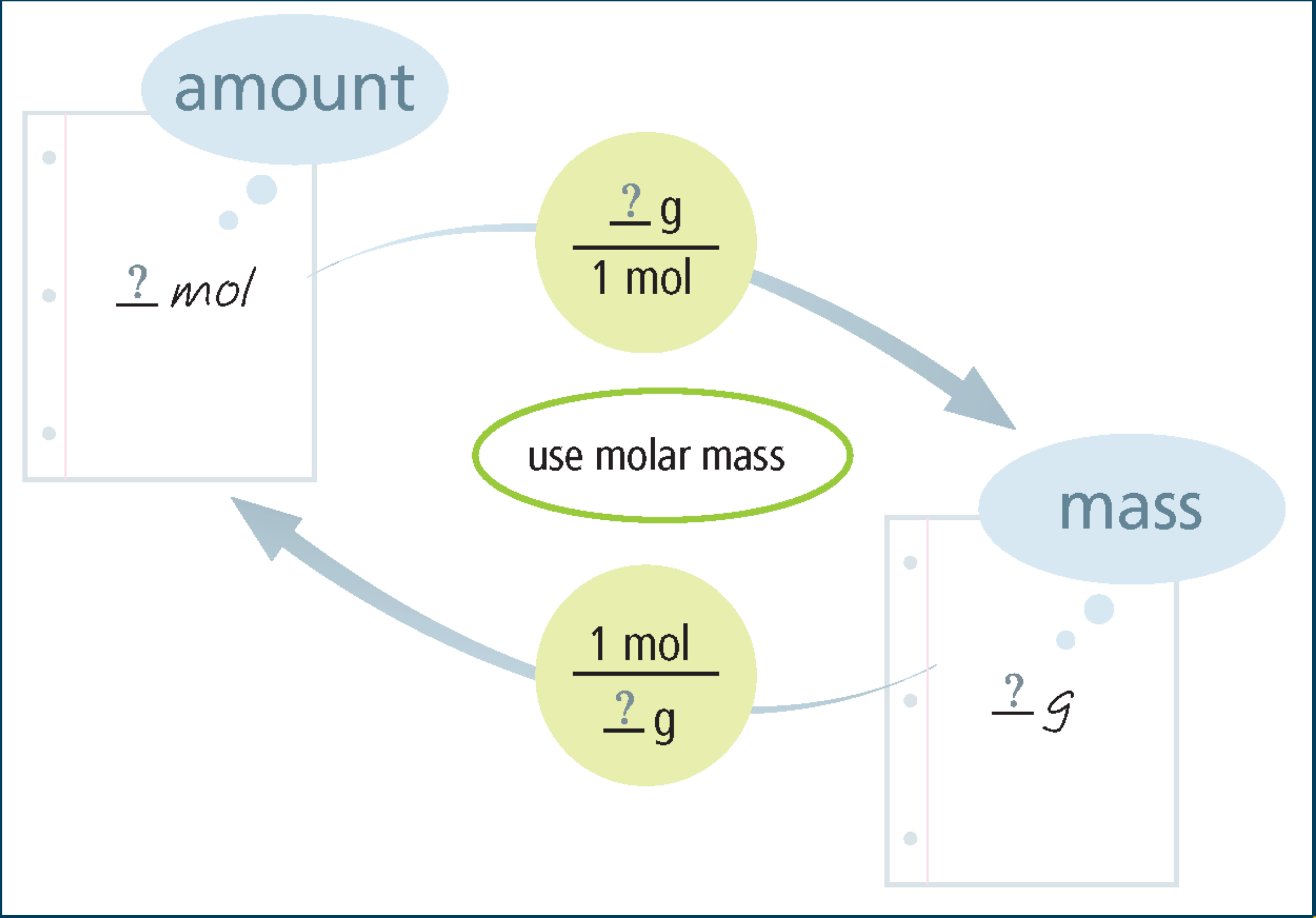
Click below to watch the Visual Concept.

Visual Concept

Solving Mole Problems



Determining the Mass from the Amount in Moles



Relating Mass to Numbers of Atoms, *continued*

Sample Problem B

What is the mass in grams of 3.50 mol of the element copper, Cu?

Relating Mass to Numbers of Atoms, *continued*

Sample Problem B Solution

Given: 3.50 mol Cu

Unknown: mass of Cu in grams

Solution: the mass of an element in grams can be calculated by multiplying the amount of the element in moles by the element's molar mass.

$$\text{moles Cu} \times \frac{\text{grams Cu}}{\text{moles Cu}} = \text{grams Cu}$$

End of Slide

Relating Mass to Numbers of Atoms, *continued*

Sample Problem B Solution, *continued*

The molar mass of copper from the periodic table is rounded to 63.55 g/mol.

$$3.50 \text{ mol Cu} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = \boxed{222 \text{ g Cu}}$$

End of Slide

Relating Mass to Numbers of Atoms, *continued*

Sample Problem C

A chemist produced 11.9 g of aluminum, Al. How many moles of aluminum were produced?

Relating Mass to Numbers of Atoms, *continued*

Sample Problem C Solution

Given: 11.9 g Al

Unknown: amount of Al in moles

Solution:

$$\text{grams Al} \times \frac{\text{moles Al}}{\text{grams Al}} = \text{moles Al}$$

The molar mass of aluminum from the periodic table is rounded to 26.98 g/mol.

$$11.9 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = \boxed{0.441 \text{ mol Al}}$$

End Of Slide

Relating Mass to Numbers of Atoms, *continued*

Sample Problem D

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

End of Slide

Relating Mass to Numbers of Atoms, *continued*

Sample Problem D Solution

Given: 3.01×10^{23} atoms of Ag

Unknown: amount of Ag in moles

Solution:

$$\text{Ag atoms} \times \frac{\text{moles Ag}}{\text{Avogadro's number of Ag atoms}} = \text{moles Ag}$$

$$3.01 \times 10^{23}\ \text{Ag atoms} \times \frac{1\ \text{mol Ag}}{6.022 \times 10^{23}\ \text{Ag atoms}} =$$

0.500 mol Ag

End Of Slide

Relating Mass to Numbers of Atoms, *continued*

Sample Problem E

What is the mass in grams of 1.20×10^8 atoms of copper, Cu?

End of Slide

Relating Mass to Numbers of Atoms, *continued*

Sample Problem E Solution

Given: 1.20×10^8 atoms of Cu

Unknown: mass of Cu in grams

Solution:

$$\text{Cu atoms} \times \frac{\text{moles Cu}}{\text{Avogadro's number of Cu atoms}} \times \frac{\text{grams Cu}}{\text{moles Cu}} = \text{grams Cu}$$

The molar mass of copper from the periodic table is rounded to 63.55 g/mol.

$$1.20 \times 10^8 \text{ Cu atoms} \times \frac{1 \text{ mol Cu}}{6.022 \times 10^{23} \text{ Cu atoms}} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} =$$

$$\boxed{1.27 \times 10^{-14} \text{ g Cu}}$$

End Of Slide

End of Chapter 3 Show

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