2 Atoms and Elements

Review Questions

2.1 Scanning tunnelling microscopy is a technique that can image, and even move, individual atoms and molecules. A scanning tunnelling microscope works by moving an extremely sharp electrode over a surface and measuring the resulting tunnelling current, the electrical current that flows between the tip of the electrode, and the surface even though the two are not in physical contact.

2.2 The first people to propose that matter was composed of small, indestructible particles were Leucippus and Democritus. These Greek philosophers theorized that matter was ultimately composed of small, indivisible particles called *atomos*. In the sixteenth century modern science began to emerge. A greater emphasis on observation brought rapid advancement as the scientific method became the established way to learn about the physical world. By the early 1800s certain observations led the English chemist John Dalton to offer convincing evidence that supported the early atomic ideas of Leucippus and Democritus. The theory that all matter is composed of atoms grew out of observations and laws. The three most important laws that led to the development and acceptance of the atomic theory were the law of conservation of mass, the law of definite proportions, and the law of multiple proportions. John Dalton explained the laws with his atomic theory.

2.3 The law of conservation of mass states the following: In a chemical reaction, matter is neither created nor destroyed. In other words, when you carry out any chemical reaction, the total mass of the substances involved in the reaction does not change.

2.4 The law of definite proportions states the following: All samples of a given compound, regardless of their source or how they were prepared, have the same proportions of their constituent elements. This means that elements composing a given compound always occur in fixed (or definite) proportions in all samples of the compound.

2.5 The law of multiple proportions states the following: When two elements (call them A and B) form two different compounds, the masses of element B that combine with 1 g of element A can be expressed as a ratio of small whole numbers. This means that when two atoms (A and B) combine to form more than one compound, the ratio of B in one compound to B in the second compound will be a small whole number.

2.6 The main ideas of John Dalton's atomic theory are as follows: 1) Each element is composed of tiny, indestructible particles called atoms. 2) All atoms of a given element have the same mass and other properties that distinguish them from the atoms of other elements. 3) Atoms combine in simple, whole number ratios to form compounds. 4) Atoms of one element cannot change into atoms of another element. They can, actually, through nuclear decay. In a chemical reaction, atoms change the way that they are bound together with other atoms to form a new substance. The law of conservation of mass is explained by the fourth idea. Since the atoms cannot change into another element, and just change how they are bound together, the total mass will remain constant. The law of constant composition is supported by idea 2 and 3. Since the atoms of a given element always have the same mass and other distinguishing properties, and they combine in simple whole number ratios, different samples of the same compound will have the same properties and the same composition. The law of multiple proportions is also supported by ideas 2 and 3 since the atoms can combine in simple whole number ratios; the ratio of the mass of B in one compound to the mass of B in a second compound will also be a small whole number.

2.7 In the late 1800s, an English physicist named J.J. Thomson performed experiments to probe the properties of cathode rays. Thomson found that these rays were actually streams of particles with the following properties: They travelled in straight lines, they were independent of the composition of the material from which they originated, and they carried a negative electrical charge. He measured the charge to mass ratio of the particles and found that the cathode ray particle was about 2000 times lighter than hydrogen.

2.8 In Millikan's oil drop experiment, oil was sprayed into fine droplets using an atomizer. The droplets were allowed to fall under the influence of gravity through a small hole into the lower portion of the apparatus where they could be viewed. During their fall, the drops would acquire electrons that had been produced by the interaction of high energy radiation with air. These charged drops interacted with two electrically charged plates within the apparatus. The negatively charged plate at the bottom of the apparatus repelled the negatively charged drops. By varying the voltage on the plates, the fall of the charged drops could be slowed, stopped, or even reversed. From the voltage required to halt the free fall of the drops, and from the masses of the drops themselves, Millikan calculated the charge of each drop. He then reasoned that, since each drop must contain an integral number of electrons, the charge of each drop must be a whole number multiple of the electron's charge. The magnitude of the charge of the electron is of tremendous importance because it determines how strongly an atom holds its electrons.

2.9 Rutherford's gold foil experiment directed positively charged α particles at an ultrathin sheet of gold foil. These particles were to act as probes of the gold atoms' structures. If the gold atoms were indeed like plum pudding—with their mass and charge spread throughout the entire volume of the atom—these speeding probes should pass right through the gold foil with minimum deflection. A majority of the particles did pass directly through the foil, but some particles were deflected, and some even bounced back. He realized that to account for the deflections, the mass and positive charge of an atom must all be concentrated in a space much smaller than the size of the atom itself.

2.10 Rutherford's nuclear model of the atom has three basic parts: 1) Most of the atom's mass and all of its positive charge are contained in a small core called the **nucleus.** 2) Most of the volume of the atom is empty space, throughout which tiny negatively charged electrons are dispersed. 3) There are as many negatively charged electrons outside the nucleus as there are positively charged particles within the nucleus, so that the atom is electrically neutral. The revolutionary part of this theory is the idea that matter, at its core, is much less uniform than it appears.

2.11 Matter appears solid because the variation in its density is on such a small scale that our eyes cannot see it.

2.12 The three subatomic particles that compose atoms are as follows:

Protons, which have a mass of 1.67262×10^{-27} kg or 1.00727 u and a relative charge of +1

Neutrons, which have a mass of 1.67493×10^{-27} kg or 1.00866 u and a relative charge of 0

Electrons, which have a mass of 0.00091×10^{-27} kg or 0.00055 u and a relative charge of −1

2.13 The number of protons in the nucleus defines the identity of an element.

2.14 The atomic number, Z, is the number of protons in an atom's nucleus. The atomic mass number (A) is the sum of the neutrons and protons in an atom.

2.15 Isotopes are atoms with the same number of protons but different numbers of neutrons. The percent natural abundance is the relative amount of each different isotope in a naturally occurring sample of a given element.

2.16 Isotopes can be symbolized as ${}^{A}_{Z}X$, where A is the mass number, Z is the atomic number, and X is the chemical symbol. A second notation is the chemical symbol (or chemical name) followed by a dash and the mass number of the isotope, such as X-A, where X is the chemical symbol or name and A is the mass number. The carbon isotope with a mass of 12 would have the symbol ${}^{12}_{6}C$ or C-12.

2.17 An ion is a charged particle. Positively charged ions are called cations. Negatively charged ions are called anions.

2.18 Atomic mass represents the average mass of the isotopes that compose that element. The average calculated atomic mass is weighted according to the natural abundance of each isotope.
Atomic mass $= \sum_n$ (fraction of isotope n) x (mass of isotope n)

2.19 In a mass spectrometer, the sample is injected into the instrument and vaporized. The vaporized atoms are then ionized by an electron beam. The electrons in the beam collide with the vaporized atoms, removing electrons from the atoms and creating positively charged ions. Charged plates with slits in them accelerate the positively charged ions into a magnetic field, which deflects them. The amount of deflection depends on the mass of the ions—lighter ions are deflected more than heavier ones. Finally, the ions strike a detector and produce an electrical signal that is recorded.

2.20 The result of the mass spectrometer is the separation of the atoms in the sample according to their mass, producing a mass spectrum. The position of each peak on the *x*-axis gives the mass of the isotope, and the intensity (indicated by the height of the peak) gives the relative abundance of that isotope.

2.21 A mole is an amount of material. It is defined as the amount of material containing 6.0221421×10^{23} particles (Avogadro's number). The numerical value of the mole is defined as being equal to the number of atoms in exactly 12 grams of pure carbon-12. It is useful for converting number of atoms to moles of atoms and moles of atoms to number of atoms.

2.22 The mass corresponding to a mole of one element is different from the mass corresponding to a mole of another element because the mass of the atom of each element is different. A mole is a specific number of atoms, so the heavier the mass of each atom, the heavier the mass of one mole of atoms.

2.23 The periodic law states the following: When elements are arranged in order of increasing mass, certain sets of properties recur periodically. Mendeleev organized all the known elements in a table consisting of a series of rows in which mass increased from left to right. The rows were arranged so that elements with similar properties were aligned in the same vertical column.

2.24 Metals are found on the left side and the middle of the periodic table. They are good conductors of heat and electricity; they can be pounded into flat sheets (malleable), they can be drawn into wires (ductile), they are often shiny, and they tend to lose electrons when they undergo chemical changes.

Nonmetals are found on the upper-right side of the periodic table. Their properties are more varied: Some are solids at room temperature, while others are liquids or gases. As a whole they tend to be poor conductors of heat and electricity and they all tend to gain electrons when they undergo chemical changes.

Metalloids lie along the zigzag diagonal line that divides metals and nonmetals. They show mixed properties. Several metalloids are also classified as semiconductors because of their intermediate and temperature-dependent electrical conductivity.

2.25 (a) Noble gases are in group 18 and are mostly unreactive. As the name implies, they are all gases in their natural state.

(b) Alkali metals are in group 1 and are all reactive metals.

(c) Alkaline earth metals are in group 2 and are also fairly reactive.

(d) Halogens are in group 17 and are very reactive nonmetals.

2.26 Main group metals tend to lose electrons, forming cations with the same number of electrons as the nearest noble gas. Main group nonmetals tend to gain electrons, forming anions with the same number of electrons as the nearest following noble gas.

Problems by Topic

The Laws of Conservation of Mass, Definite Proportions, and Multiple Proportions

2.27 **Given:** 1.50 g hydrogen; 11.9 g oxygen **Find:** grams water vapour
Conceptual Plan: total mass reactants = total mass products
Solution: Mass of reactants = 1.50 g hydrogen + 11.9 g oxygen = 13.4 grams
Mass of products = mass of reactants = 13.4 grams water vapour.
Check: According to the law of conservation of mass, matter is not created or destroyed in a chemical reaction, so, since water vapour is the only product, the masses of hydrogen and oxygen must combine to form the mass of water vapour.

2.28 **Given:** 21 kg gasoline; 84 kg oxygen **Find:** mass of carbon dioxide and water
Conceptual Plan: total mass reactants = total mass products
Solution: Mass of reactants = 21 kg gasoline + 84 kg oxygen = 105 kg mass
Mass of products = mass of reactants = 105 kg of mass of carbon dioxide and water.
Check: According to the law of conservation of mass, matter is not created or destroyed in a chemical reaction, so, since carbon dioxide and water are the only products, the masses of gasoline and oxygen must combine to form the mass of carbon dioxide and water.

2.29 **Given:** sample 1: 38.9 g carbon, 448 g chlorine; sample 2: 14.8 g carbon, 134 g chlorine
Find: are results consistent with definite proportions?
Conceptual Plan: determine mass ratio of sample 1 and 2 and compare

$$\frac{\text{mass of chlorine}}{\text{mass of carbon}}$$

Solution: Sample 1: $\dfrac{448\text{ g chlorine}}{38.9\text{ g carbon}} = 11.5$ Sample 2: $\dfrac{134\text{ g chlorine}}{14.8\text{ g carbon}} = 9.05$

Results are not consistent with the law of definite proportions because the ratio of chlorine to carbon is not the same.
Check: According to the law of definite proportions, the mass ratio of one element to another is the same for all samples of the compound.

2.30 **Given:** sample 1: 6.98 grams sodium, 10.7 grams chlorine; sample 2: 11.2 g sodium, 17.3 grams chlorine
Find: are results consistent with definite proportions?
Conceptual Plan: determine mass ratio of sample 1 and 2 and compare

$$\frac{\text{mass of chlorine}}{\text{mass of sodium}}$$

Solution: Sample 1: $\dfrac{10.7\text{ g chlorine}}{6.98\text{ g sodium}} = 1.53$ Sample 2: $\dfrac{17.3\text{ g chlorine}}{11.2\text{ g sodium}} = 1.54$

Results are consistent with the law of definite proportions.
Check: According to the law of definite proportions, the mass ratio of one element to another is the same for all samples of the compound.

2.31 **Given:** mass ratio sodium to fluorine = 1.21:1; sample = 28.8 g sodium **Find:** g fluorine
Conceptual Plan: g sodium $\rightarrow$ g fluorine

$$\frac{\text{mass of fluorine}}{\text{mass of sodium}}$$

Solution: $28.8\,\cancel{\text{g sodium}} \times \dfrac{1\text{ g fluorine}}{1.21\,\cancel{\text{g sodium}}} = 23.8\text{ g fluorine}$

Check: The units of the answer (g fluorine) are correct. The magnitude of the answer is reasonable since it is less than the grams of sodium.

2.32 **Given:** sample 1: 1.65 kg magnesium, 2.57 kg fluorine; sample 2: 1.32 kg magnesium
Find: g fluorine in sample 2
Conceptual Plan: mass magnesium and mass fluorine $\xrightarrow{\frac{\text{mass of fluorine}}{\text{mass of magnesium}}}$ mass ratio → mass fluorine(kg) $\xrightarrow{\frac{1000\text{ g}}{\text{kg}}}$ mass fluorine(g)

Solution: $\text{mass ratio} = \dfrac{2.57\text{ kg fluorine}}{1.65\text{ kg magnesium}} = \dfrac{1.56\text{ kg fluorine}}{1.00\text{ kg magnesium}}$

$$1.32\,\cancel{\text{kg magnesium}} \times \frac{1.56\text{ kg fluorine}}{1.00\,\cancel{\text{kg magnesium}}} \times \frac{1000\text{ g}}{\text{kg}} = 2.06 \times 10^3\text{ g fluorine}$$

Check: The units of the answer (g fluorine) are correct. The magnitude of the answer is reasonable since it is greater than the mass of magnesium and the ratio is greater than 1.

2.33 **Given:** 1 gram osmium: sample 1 = 0.168 g oxygen; sample 2 = 0.3369 g oxygen
Find: are results consistent with multiple proportions?
Conceptual Plan: determine mass ratio of oxygen

$$\frac{\text{mass of oxygen sample 2}}{\text{mass of oxygen sample 1}}$$

Solution: $\dfrac{0.3369\text{ g oxygen}}{0.168\text{ g oxygen}} = 2.00$ Ratio is a small whole number. Results are consistent with multiple proportions
Check: According to the law of multiple proportions, when two elements form two different compounds, the masses of element B that combine with 1 g of element A can be expressed as a ratio of small whole numbers.

2.34 **Given:** 1 g palladium: compound A: 0.603 g S; compound B: 0.301 g S; compound C: 0.151 g S
Find: are results consistent with multiple proportions?
Conceptual Plan: determine mass ratio of sulfur in the three compounds

$$\frac{\text{mass of sulfur sample A}}{\text{mass of sulfur sample B}} \quad \frac{\text{mass of sulfur sample A}}{\text{mass of sulfur sample C}} \quad \frac{\text{mass of sulfur sample B}}{\text{mass of sulfur sample C}}$$

Solution: $\dfrac{0.603\text{ g S in compound A}}{0.301\text{ g S in compound B}} = 2.00 \qquad \dfrac{0.603\text{ g S in compound A}}{0.151\text{ g S in compound C}} = 3.99 \sim 4$

$$\frac{0.301\text{ g S in compound B}}{0.151\text{ g S in compound C}} = 1.99 \sim 2$$

Ratio of each is a small whole number. Results are consistent with multiple proportions.
Check: According to the law of multiple proportions, when two elements form two different compounds, the masses of element B that combine with 1 g of element A can be expressed as a ratio of small whole numbers.

2.35 **Given:** sulfur dioxide = 3.49 g oxygen and 3.50 g sulfur; sulfur trioxide = 6.75 g oxygen and 4.50 g sulfur
Find: mass oxygen per g S for each compound and then determine the mass ratio of oxygen

$$\frac{\text{mass of oxygen in sulfur dioxide}}{\text{mass of sulfur in sulfur dioxide}} \quad \frac{\text{mass of oxygen in sulfur trioxide}}{\text{mass of sulfur in sulfur trioxide}} \quad \frac{\text{mass of oxyen in sulfur trioxide}}{\text{mass of oxyen in sulfur dioxide}}$$

Solution: $\text{sulfur dioxide} = \dfrac{3.49\text{ g oxygen}}{3.50\text{ g sulfur}} = \dfrac{0.997\text{ g oxygen}}{1\text{ g sulfur}}$ $\text{sulfur trioxide} = \dfrac{6.75\text{ g oxygen}}{4.50\text{ g sulfur}} = \dfrac{1.50\text{ g oxygen}}{1\text{ g sulfur}}$

$$\frac{1.50\text{ g oxygen in sulfur trioxide}}{0.997\text{ g oxygen in sulfur dioxide}} = \frac{1.50}{1} = \frac{3}{2}$$

Ratio is in small whole numbers and is consistent with multiple proportions.
Check: According to the law of multiple proportions, when two elements form two different compounds, the masses of element B that combine with 1 g of element A can be expressed as a ratio of small whole numbers.

2.36 **Given:** sulfur hexafluoride = 4.45 g fluorine and 1.25 g sulfur; sulfur tetrafluoride = 4.43 g fluorine and 1.87 g sulfur
Find: mass fluorine per g S for each compound and then determine the mass ratio of fluorine

$$\frac{\text{mass of fluorine in sulfur hexafluoride}}{\text{mass of sulfur in sulfur hexafluoride}} \quad \frac{\text{mass of fluorine in sulfur tetrafluoride}}{\text{mass of sulfur in sulfur tetrafluoride}} \quad \frac{\text{mass of fluorine in sulfur hexafluoride}}{\text{mass of fluorine in sulfur tetrafluoride}}$$

Solution: $\text{sulfur hexafluoride} = \dfrac{4.45\text{ g fluorine}}{1.25\text{ g sulfur}} = \dfrac{3.56\text{ g fluorine}}{1\text{ g sulfur}}$

$\text{sulfur tetrafluoride} = \dfrac{4.43\text{ g fluorine}}{1.87\text{ g sulfur}} = \dfrac{2.3\underline{6}9\text{ g fluorine}}{1\text{ g sulfur}}$

$\dfrac{3.56\text{ g fluorine in sulfur hexafluoride}}{2.3\underline{6}9\text{ g fluorine in sulfur tetrafluoride}} = \dfrac{1.50}{1} = \dfrac{3}{2}$

Ratio is in small whole numbers and is consistent with multiple proportions.

Check: According to the law of multiple proportions, when two elements form two different compounds, the masses of element B that combine with 1 g of element A can be expressed as a ratio of small whole numbers.

Atomic Theory, Nuclear Theory, and Subatomic Particles

2.37 **Given:** drop A = -6.9×10^{-19} C; drop B = -9.2×10^{-19} C; drop C = -11.5×10^{-19} C; drop D = -4.6×10^{-19} C
Find: the charge on a single electron
Conceptual Plan: determine the ratio of charge for each set of drops

$\dfrac{\text{charge on drop 1}}{\text{charge on drop 2}}$

Solution: $\dfrac{-6.9 \times 10^{-19}\text{C drop A}}{-4.6 \times 10^{-19}\text{ C drop D}} = 1.5 \quad \dfrac{-9.2 \times 10^{-19}\text{C drop B}}{-4.6 \times 10^{-19}\text{ C drop D}} = 2 \quad \dfrac{-11.5 \times 10^{-19}\text{C drop C}}{-4.6 \times 10^{-19}\text{ C drop D}} = 2.5$

The ratios obtained are not whole numbers, but can be converted to whole numbers by multiplying by 2.

Therefore, the charge on the electron has to be 1/2 the smallest value experimentally obtained. The charge on the electron = -2.3×10^{-19} C.

Check: The units of the answer (Coulombs) are correct. The magnitude of the answer is reasonable since all the values experimentally obtained are integer multiples of -2.3×10^{-19}.

2.38 **Given:** $m_{drop} = 5.13 \times 10^{-15}$ kg, $V = 350$ V, $g = 9.807$ m s^{-2}, $d = 1$ cm
Find: n_e, number of excess electrons on oil drop
Other: $e_c = 1.60 \times 10^{-19}$ C
Conceptual Plan: $q = mgd/V$. $q, e_c \rightarrow n_e$

Solution: $q = mgd/V$

$$q = \frac{-(5.13 \times 10^{-15}\text{ kg})(9.807\text{ m s}^{-2})(0.01\text{ m})}{350\text{ kg m}^2\text{ s}^{-2}\text{C}^{-1}} = -1.4\underline{3}7426 \times 10^{-18}\text{ C}$$

$$\frac{-1.4\underline{3}7426 \times 10^{-18}\text{ C}}{\text{drop}} \times \frac{1\text{ electron}}{-1.60 \times 10^{-19}\text{ C}} = 8.98\text{ electrons/drop} = 9\text{ electrons/drop}$$

Check: In his oil drop experiment, Millikin discovered that the charge measured on each oil droplet was an integer multiple of 1.6×10^{-9} C, which is the fundamental charge of an electron. Therefore, 9 electrons per oil droplet makes sense.

2.39 **Given:** charge on body = -15 μC **Find:** number of electrons, mass of the electrons
Conceptual Plan: μC → C → number of electrons → mass of electrons

$\dfrac{1\text{ C}}{10^6\,\mu\text{C}} \quad \dfrac{1\text{ electron}}{-1.60 \times 10^{-19}\text{C}} \quad \dfrac{9.10 \times 10^{-28}\text{ g}}{1\text{ electron}}$

Solution: $-15\,\cancel{\mu\text{C}} \times \dfrac{1\,\cancel{\text{C}}}{10^6\,\cancel{\mu\text{C}}} \times \dfrac{1\text{ electron}}{-1.60 \times 10^{-19}\,\cancel{\text{C}}} = 9.\underline{3}75 \times 10^{13}\text{ electrons} = 9.4 \times 10^{13}\text{ electrons}$

$$9.\underline{3}75 \times 10^{13}\,\cancel{\text{electrons}} \times \frac{9.10 \times 10^{-28}\text{ g}}{1\,\cancel{\text{electron}}} = 8.5 \times 10^{-14}\text{ g} = 8.5 \times 10^{-17}\text{ kg}$$

Check: The units of the answers (number of electrons and grams) are correct. The magnitude of the answers is reasonable since the charge on an electron and the mass of an electron are very small.

2.40 **Given:** charge = –1.0 C **Find:** number of electrons, mass of the electrons

Conceptual Plan: C $\rightarrow$ number of electrons $\rightarrow$ mass of electrons

$$\frac{1 \text{ electron}}{-1.60 \times 10^{-19} \text{C}} \qquad \frac{9.10 \times 10^{-28} \text{ g}}{1 \text{ electron}}$$

Solution: $-1.0 \,\cancel{\text{C}} \times \dfrac{1 \text{ electron}}{-1.60 \times 10^{-19} \,\cancel{\text{C}}} = 6.\underline{2}5 \times 10^{18} \text{ electrons} = 6.3 \times 10^{18} \text{ electrons}$

$$6.\underline{2}5 \times 10^{18} \,\cancel{\text{electrons}} \times \frac{9.10 \times 10^{-28} \text{ g}}{1 \,\cancel{\text{electron}}} = 5.7 \times 10^{-9} \text{ g} = 5.7 \times 10^{-12} \text{ kg}$$

Check: The units of the answers (number of electrons and grams) are correct. The magnitude of the answers is reasonable since the charge on an electron and the mass of an electron are very small.

2.41 **Given:** mass of proton **Find:** number of electron in equal mass

Conceptual Plan: mass of protons $\rightarrow$ number of electrons

$$\frac{1.67262 \times 10^{-27} \text{ kg}}{1 \text{ proton}} \qquad \frac{1 \text{ electron}}{9.10938 \times 10^{-31} \text{ kg}}$$

Solution: $1.67262 \times 10^{-27} \,\cancel{\text{kg}} \times \dfrac{1 \text{ electron}}{9.10938 \times 10^{-31} \,\cancel{\text{kg}}} = 1.83615 \times 10^{3} \text{ electrons} = 1836 \text{ electrons}$

Check: The units of the answer (electrons) are correct. The magnitude of the answer is reasonable since the mass of the electron is much less than the mass of the proton.

2.42 **Given:** helium nucleus **Find:** number of electrons in equal mass

Conceptual Plan:

 # protons $\rightarrow$ mass of protons and # neutrons $\rightarrow$ mass of neutrons $\rightarrow$ total mass $\rightarrow$ number of electrons

$$\frac{1.67262 \times 10^{-27} \text{ kg}}{1 \text{ proton}} \qquad \frac{1.67493 \times 10^{-27} \text{ kg}}{1 \text{ neutron}} \qquad \text{mass protons} + \text{mass neutrons} \qquad \frac{1 \text{ electron}}{0.00091 \times 10^{-27} \text{ kg}}$$

Solution: $2 \,\cancel{\text{protons}} \times \dfrac{1.67262 \times 10^{-27} \text{ kg}}{\cancel{\text{proton}}} = 3.34524 \times 10^{-27} \text{ kg}$

$$2 \,\cancel{\text{neutrons}} \times \frac{1.67493 \times 10^{-27} \text{ kg}}{\cancel{\text{neutron}}} = 3.34986 \times 10^{-27} \text{ kg}$$

$$(3.34524 \times 10^{-27} \,\cancel{\text{kg}} + 3.34986 \times 10^{-27} \,\cancel{\text{kg}}) \times \frac{1 \text{ electron}}{9.10938 \times 10^{-31} \,\cancel{\text{kg}}} = 7.34968 \times 10^{3} \text{ electrons} = 7350 \text{ electrons}$$

Check: The units of the answer (electrons) are correct. The magnitude of the answer is reasonable since the mass of the electrons is much less than the mass of the proton and neutron.

Isotopes and Ions

2.43 For each of the isotopes determine Z (the number of protons) from the periodic table and determine A (protons + neutrons). Then, write the symbol in the form $^{A}_{Z}X$.

(a) The copper isotope with 34 neutrons: Z = 29; A = 29 + 34 = 63; $^{63}_{29}Cu$

(b) The copper isotope with 36 neutrons: Z = 29; A = 29 + 36 = 65; $^{65}_{29}Cu$

(c) The potassium isotope with 21 neutrons: Z = 19; A = 19 + 21 = 40; $^{40}_{19}K$

(d) The argon isotope with 22 neutrons: Z = 18; A = 18 + 22 = 40; $^{40}_{18}Ar$

2.44 For each of the isotopes determine Z (the number of protons) from the periodic table and determine A (protons + neutrons). Then, write the symbol in the form X-A.

(a) The silver isotope with 60 neutrons: Z = 47; A = 47 + 60 = 107; Ag-107

(b) The silver isotope with 62 neutrons: Z = 47; A = 47 + 62 = 109; Ag-109

(c) The uranium isotope with 146 neutrons: Z = 92; A = 92 + 146 = 238; U-238

(d) The hydrogen isotope with 1 neutron: Z = 1: A = 1 + 1 = 2; H-2

2.45 (a) ${}^{14}_{7}N$: Z = 7 ; A = 14; protons = Z = 7; neutrons = A – Z = 14 – 7 = 7

(b) ${}^{23}_{11}Na$: Z = 11; A = 23; protons = Z = 11; neutrons = A – Z = 23 – 11 = 12

(c) ${}^{222}_{86}Rn$: Z = 86; A = 222; protons = Z = 86; neutrons = A – Z = 222 – 86 = 136

(d) ${}^{208}_{82}Pb$: Z = 82; A = 208; protons = Z = 82; neutrons = A – Z = 208 – 82 = 126

2.46 (a) ${}^{40}_{19}K$: Z = 19; A = 40; protons = Z = 19; neutrons = A – Z = 40 – 19 = 21

(b) ${}^{226}_{88}Ra$: Z = 88; A = 226; protons = Z = 88; neutrons = A – Z = 226 – 88 = 138

(c) ${}^{99}_{43}Tc$: Z = 43; A = 99; protons = Z = 43; neutrons = A – Z = 99 – 43 = 56

(d) ${}^{33}_{15}P$: Z = 15; A = 33; protons = Z = 15; neutrons = A – Z = 33 – 15 = 18

2.47 Carbon – 14: A = 14, Z = 6: ${}^{14}_{6}C$ # protons = Z = 6 # neutrons = A – Z = 14 – 6 = 8

2.48 Uranium – 235: A = 235, Z = 92: ${}^{235}_{92}U$ # protons = Z = 92 # neutrons = A – Z = 235 – 92 = 143

2.49 In a neutral atom the number of protons = the number of electrons = Z. For an ion, electrons are lost (cations) or gained (anions).

(a) Ni^{2+}: Z = 28 = protons; Z – 2 = 26 = electrons

(b) S^{2-}: Z = 16 = protons; Z + 2 = 18 = electrons

(c) Br^{-}: Z = 35 = protons; Z + 1 = 36 = electrons

(d) Cr^{3+}: Z = 24 = protons; Z – 3 = 21 = electrons

2.50 In a neutral atom the number of protons = the number of electrons = Z. For an ion, electrons are lost (cations) or gained (anions).

(a) Al^{3+}: Z = 13 = protons; Z – 3 = 10 = electrons

(b) Se^{2-}: Z = 34 = protons; Z + 2 = 36 = electrons

(c) Ga^{3+}: Z = 31 = protons; Z – 3 = 28 = electrons

(d) Sr^{2+}: Z = 38 = protons; Z – 2 = 36 = electrons

2.51 Main group metal atoms will lose electrons to form a cation with the same number of electrons as the nearest, previous noble gas. Atoms in period 4 and higher lose electrons to form the same ion as the element at the top of the group.

Nonmetal atoms will gain electrons to form an anion with the same number of electrons as the nearest noble gas.

Symbol	**Ion Formed**	**Number of Electrons in Ion**	**Number of Protons in Ion**
Ca	Ca^{2+}	**18**	**20**
Be	Be^{2+}	2	**4**
Se	$\mathbf{Se^{2-}}$	**36**	34
In	In^{3+}	**46**	**49**

2.52 Main group metal atoms will lose electrons to form a cation with the same number of electrons as the nearest, previous noble gas.

Nonmetal atoms will gain electrons to form an anion with the same number of electrons as the nearest noble gas.

Symbol	Ion Formed	Number of Electrons in Ion	Number of Protons in Ion
Cl	**Cl^-**	**18**	17
Te	**Te^{2-}**	54	**52**
Br	Br^-	**36**	**35**
Sr	Sr^{2+}	**36**	**38**

Atomic Mass

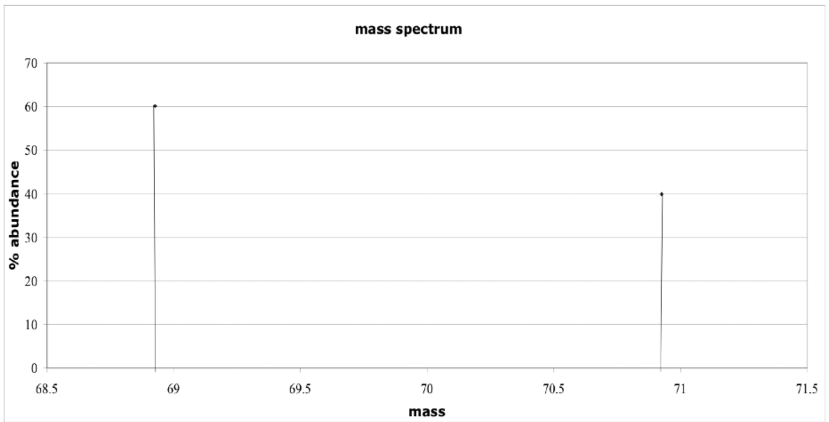
2.53 **Given:** Ga-69; mass = 68.92558 u; 60.108%: Ga-71; mass = 70.92470 u; 39.892% **Find:** atomic mass Ga
Conceptual Plan: % abundance → fraction and then find atomic mass

$$\frac{\%\text{ abundance}}{100} \quad \text{Atomic mass} = \sum_n (\text{fraction of isotope } n) \times (\text{mass of isotope } n)$$

Solution: Fraction Ga-69 $= \dfrac{60.108}{100} = 0.60108$ Fraction Ga-71 $= \dfrac{39.892}{100} = 0.39892$

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

$$= 0.60108(68.92588\text{ u}) + 0.39892(70.92470\text{ u}) = 69.723\text{ u}$$



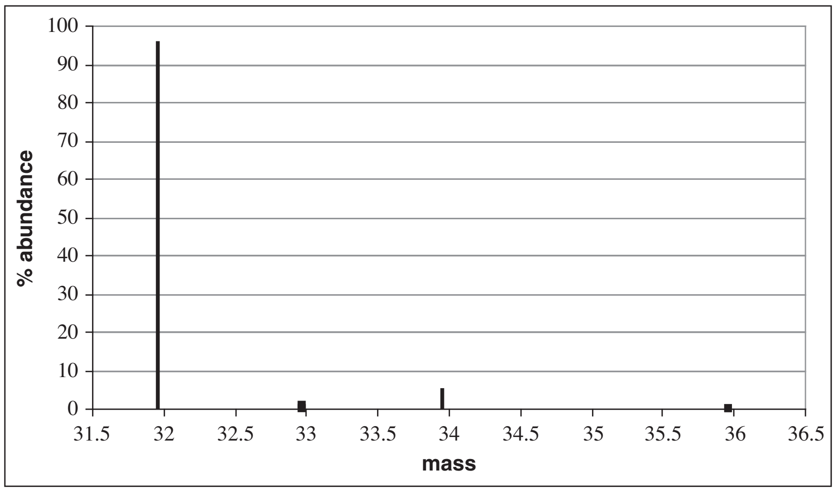
Check: Units of the answer (u) are correct. The magnitude of the answer is reasonable because it lies between 68.92588 u and 70.92470 u and is closer to 68.92588, which has the higher % abundance. The mass spectrum is reasonable because it has two mass lines corresponding to the two isotopes, and the line at 68.92588 is about 1.5 times larger than the line at 70.92470.

2.54 **Given:** sulfur isotopes, masses in u and % abundances given in the table in the problem
Find: atomic mass of sulfur
Conceptual Plan: The atomic mass of an element is the weighted average of its constituent isotopes. That is:

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

Solution: Atomic mass $= (0.9499 \times 31.9721\text{ u}) + (0.0075 \times 32.9715\text{ u}) + (0.0425 \times 33.9679\text{ u})$
$+ (0.0001 \times 35.9671\text{ u})$
$= 32.06\text{ u}$

Mass spectrum is shown below:



Check: The unit is correct and has the correct number of significant figures. The answer itself is in agreement with published value

2.55 Fluorine exists only as F-19, so the mass spectrum of fluorine exhibits just one line at 18.998 u. Chlorine has two isotopes, Cl-35 and Cl-37, and the mass of 35.45 u is the weighted average of these two isotopes, so there is no line at 35.45 u.

2.56 Copper has no isotope with a mass of 63.546 u. Since the mass of the isotope comes primarily from the sum of the protons and neutrons, the mass of the isotope has to have a value close to a whole number. Copper must be composed of two or more isotopes, one with a mass less than 63.546 u and one with a mass greater than 63.546 u.

2.57 **Given:** isotope – 1 mass = 120.9038 u, 57.4%; isotope – 2 mass = 122.9042 u
Find: atomic mass of the element and identify the element
Conceptual Plan:
% abundance isotope 2 → and then % abundance → fraction and then find atomic mass

$$100\% - \%\text{ abundance isotope 1} \qquad \frac{\%\text{ abundance}}{100} \qquad \text{Atomic mass} = \sum_n (\text{fraction of isotope } n) \times (\text{mass of isotope } n)$$

Solution: 100.0% – 57.4% isotope 1 = 42.6% isotope 2

$$\text{Fraction isotope 1} = \frac{57.4}{100} = 0.574 \qquad \text{Fraction isotope 2} = \frac{42.6}{100} = 0.426$$

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

$$= 0.574(120.9038\text{ u}) + 0.426(122.9042\text{ u}) = 121.8\text{ u}$$

From the periodic table, Sb has a mass of 121.757 u, so it is the closest mass and the element is antimony.
Check: The units of the answer (u) are correct. The magnitude of the answer is reasonable because it lies between 120.9038 and 122.9042 and is slightly less than halfway between the two values because the lower value has a slightly greater abundance.

2.58 **Given:** isotope – 1 mass = 135.90714 u, 50.19%; isotope – 2 mass = 137.90599 u, 0.25%; isotope – 3 mass = 139.90543 u, 88.43%; isotope – 4, mass = 141.90924, 11.13%
Find: atomic mass of the element and identify the element
Conceptual Plan: % abundance → fraction and then find atomic mass

$$\frac{\%\text{ abundance}}{100} \qquad \text{Atomic mass} = \sum_n (\text{fraction of isotope } n) \times (\text{mass of isotope } n)$$

Solution: Fraction isotope 1 $= \frac{0.19}{100} = 0.0019$ Fraction isotope 2 $= \frac{0.25}{100} = 0.0025$

Fraction isotope 3 $= \frac{88.43}{100} = 0.8843$ Fraction isotope 4 $= \frac{11.13}{100} = 0.1113$

Atomic mass $= \sum_n$(fraction of isotope n) x (mass of isotope n)

$= 0.0019(135.90714 \text{ u}) + 0.0025(137.90599 \text{ u}) + 0.8843(139.90543 \text{ u}) + 0.1113(141.90924 \text{ u})$
$= 140.1 \text{ u}$

From the periodic table, Ce has a mass of 140.12 u, so it is the closest mass and the element is cerium.
Check: The units of the answer (u) are correct. The magnitude of the answer is reasonable because it lies between 135.90714 and 141.90924 and is closer to the higher value because isotopes 3 and 4 make up most of the mass of the element with a combined abundance of 99.56%.

2.59 **Given:** Br-81; mass = 80.9163 u; 49.31%: atomic mass Br = 79.904 u **Find:** mass and abundance
Conceptual Plan: % abundance Br-79 → then % abundance → fraction → mass Br-79

100% − %Br-81 $\frac{\text{\% abundance}}{100}$ Atomic mass $= \sum_n$(fraction of isotope n) x (mass of isotope n)

Solution: 100.00% – 49.31% = 50.69%

Fraction Br-79 $= \frac{50.69}{100} = 0.5069$ Fraction Br-81 $= \frac{49.31}{100} = 0.4931$

Let X be the mass of Br-79

Atomic mass $= \sum_n$(fraction of isotope n) x (mass of isotope n)

$79.904 \text{ u} = 0.5069(\text{X u}) + 0.4931(80.9163 \text{ u})$

$\text{X} = 78.92 \text{ u} =$ mass Br-79

Check: The units of the answer (u) are correct. The magnitude of the answer is reasonable because it is less than the mass of the atom, and the second isotope (Br-81) has a mass greater than the mass of the atom.

2.60 **Given:** Si-28; mass = 27.9769 u; 92.2%: Si-29; mass = 28.9765 u; 4.67%: atomic mass Si = 28.09 u
Find: mass and abundance Si-30
Conceptual Plan: % abundance Si-30 → then % abundance → fraction → mass Si-30

100% − %Si-30 $\frac{\text{\% abundance}}{100}$ Atomic mass $= \sum_n$(fraction of isotope n) x (mass of isotope n)

Solution: 100.00% – 92.2% – 4.67% = 3.1% Si-30

Fraction Si-28 $= \frac{92.2}{100} = 0.922$ Fraction Si-29 $= \frac{4.67}{100} = 0.0467$ Fraction Si-30 $= \frac{3.1}{100} = 0.031$

Let X be the mass of Si-30

Atomic mass $= \sum_n$(fraction of isotope n) x (mass of isotope n)

$28.09 \text{ u} = 0.922(27.9769 \text{ u}) + 0.0467(28.9765 \text{ u}) + 0.031(\text{X})$
$\text{X} = 30 \text{ u} =$ mass Si-30

Check: The units of the answer (u) are correct. The magnitude of the answer is reasonable because it is greater than the mass of the second isotope (Si-29), which has a mass greater than the mass of the atom.

The Mole Concept

2.61 **Given:** 3.8 mol sulfur **Find:** atoms of sulfur
Conceptual Plan: mol S → atoms S

$\frac{6.022 \times 10^{23} \text{ atoms}}{\text{mol}}$

Solution: $3.8 \cancel{\text{mol S}} \times \frac{6.022 \times 10^{23} \text{ atoms S}}{\cancel{\text{mol S}}} = 2.3 \times 10^{24}$ atoms S

Check: The units of the answer (atoms S) are correct. The magnitude of the answer is reasonable since there is more than 1 mole of material present.

2.62 **Given:** 5.8×10^{24} aluminum atoms **Find:** mol Al

Conceptual Plan: atoms Al → mol Al

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

Solution: $5.8 \times 10^{24} \cancel{\text{atoms Al}} \times \dfrac{1 \text{ mol Al}}{6.022 \times 10^{23} \cancel{\text{atoms Al}}} = 9.6 \text{ mol Al}$

Check: The units of the answer (mol Al) are correct. The magnitude of the answer is reasonable since there is greater than Avogadro's number of atoms present.

2.63 (a) **Given:** 11.8 g Ar **Find:** mol Ar

Conceptual Plan: g Ar → mol Ar

$$\frac{1 \text{ mol Ar}}{39.95 \text{ g Ar}}$$

Solution: $11.8 \cancel{\text{g Ar}} \times \dfrac{1 \text{ mol Ar}}{39.95 \cancel{\text{g Ar}}} = 0.295 \text{ mol Ar}$

Check: The units of the answer (mol Ar) are correct. The magnitude of the answer is reasonable since there is less than the mass of 1 mol present.

(b) **Given:** 3.55 g Zn **Find:** mol Zn

Conceptual Plan: g Zn → mol Zn

$$\frac{1 \text{ mol Zn}}{65.41 \text{ g Zn}}$$

Solution: $3.55 \cancel{\text{g Zn}} \times \dfrac{1 \text{ mol Zn}}{65.41 \cancel{\text{g Zn}}} = 0.0543 \text{ mol Zn}$

Check: The units of the answer (mol Zn) are correct. The magnitude of the answer is reasonable since there is less than the mass of 1 mol present.

(c) **Given:** 26.1 g Ta **Find:** mol Ta

Conceptual Plan: g Ta → mol Ta

$$\frac{1 \text{ mol Ta}}{180.95 \text{ g Ta}}$$

Solution: $26.1 \cancel{\text{g Ta}} \times \dfrac{1 \text{ mol Ta}}{180.95 \cancel{\text{g Ta}}} = 0.144 \text{ mol Ta}$

Check: The units of the answer (mol Ta) are correct. The magnitude of the answer is reasonable since there is less than the mass of 1 mol present.

(d) **Given:** 0.211 g Li **Find:** mol Li

Conceptual Plan: g Li → mol Li

$$\frac{1 \text{ mol Li}}{6.941 \text{ g Li}}$$

Solution: $0.211 \cancel{\text{g Li}} \times \dfrac{1 \text{ mol Li}}{6.941 \cancel{\text{g Li}}} = 0.0304 \text{ mol Li}$

Check: The units of the answer (mol Li) are correct. The magnitude of the answer is reasonable since there is less than the mass of 1 mol present.

2.64 (a) **Given:** 2.3×10^{-3} mol Sb **Find:** grams Sb

Conceptual Plan: mol Sb → g Sb

$$\frac{121.76 \text{ g Sb}}{1 \text{ mol Sb}}$$

Solution: $2.3 \times 10^{-3} \cancel{\text{mol Sb}} \times \dfrac{121.76 \text{ g Sb}}{1 \cancel{\text{mol Sb}}} = 0.28 \text{ grams Sb}$

Check: The units of the answer (grams Sb) are correct. The magnitude of the answer is reasonable since there is less than 1 mol of Sb present.

(b) **Given:** 0.0355 mol Ba **Find:** grams Ba

Conceptual Plan: mol Ba → g Ba

$$\frac{137.33 \text{ g Ba}}{1 \text{ mol Ba}}$$

Solution: $0.0355 \cancel{\text{mol Ba}} \times \dfrac{137.33 \text{ g Ba}}{1 \cancel{\text{mol Ba}}} = 4.88 \text{ grams Ba}$

Check: The units of the answer (grams Ba) are correct. The magnitude of the answer is reasonable since there is less than 1 mol of Ba present.

(c) **Given:** 43.9 mol Xe **Find:** grams Xe
Conceptual Plan: mol Xe → g Xe

$$\frac{131.29 \text{ g Xe}}{1 \text{ mol Xe}}$$

Solution: $43.9 \cancel{\text{mol Xe}} \times \frac{131.29 \text{ g Xe}}{1 \cancel{\text{mol Xe}}} = 5.76 \times 10^3$ grams Xe

Check: The units of the answer (grams Xe) are correct. The magnitude of the answer is reasonable since there is much more than 1 mol of Xe present.

(d) **Given:** 1.3 mol W **Find:** grams W
Conceptual Plan: mol W → g W

$$\frac{183.84 \text{ g W}}{1 \text{ mol W}}$$

Solution: $1.3 \cancel{\text{mol W}} \times \frac{183.84 \text{ g W}}{1 \cancel{\text{mol W}}} = 2.4 \times 10^2$ grams W

Check: The units of the answer (grams W) are correct. The magnitude of the answer is reasonable since there is slightly over 1 mol of W present.

2.65 **Given:** 3.78 g silver **Find:** atoms Ag
Conceptual Plan: g Ag → mol Ag → atoms Ag

$$\frac{1 \text{ mol Ag}}{107.87 \text{ g Ag}} \qquad \frac{6.022 \times 10^{23} \text{ atoms}}{\text{mol}}$$

Solution: $3.78 \cancel{\text{g Ag}} \times \frac{1 \cancel{\text{mol Ag}}}{107.87 \cancel{\text{g Ag}}} \times \frac{6.022 \times 10^{23} \text{ atoms Ag}}{1 \cancel{\text{mol Ag}}} = 2.11 \times 10^{22}$ atoms Ag

Check: The units of the answer (atoms Ag) are correct. The magnitude of the answer is reasonable since there is less than the mass of 1 mol of Ag present.

2.66 **Given:** 4.91×10^{21} Pt atoms **Find:** g Pt
Conceptual Plan: atoms Pt → mol Pt → g Pt

$$\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}} \qquad \frac{195.08 \text{ g Pt}}{1 \text{ mol Pt}}$$

Solution: $4.91 \times 10^{21} \cancel{\text{atoms Pt}} \times \frac{1 \cancel{\text{mol Pt}}}{6.022 \times 10^{23} \cancel{\text{atoms Pt}}} \times \frac{195.08 \text{ g Pt}}{1 \cancel{\text{mol Pt}}} = 1.59$ g Pt

Check: The units of the answer (g Pt) are correct. The magnitude of the answer is reasonable since there is less than 1 mol of Pt atoms present.

2.67 (a) **Given:** 5.18 g P **Find:** atoms P
Conceptual Plan: g P → mol P → atoms P

$$\frac{1 \text{ mol P}}{30.97 \text{ g P}} \qquad \frac{6.022 \times 10^{23} \text{ atoms}}{\text{mol}}$$

Solution: $5.18 \cancel{\text{g P}} \times \frac{1 \cancel{\text{mol P}}}{30.97 \cancel{\text{g P}}} \times \frac{6.022 \times 10^{23} \text{ atoms P}}{1 \cancel{\text{mol P}}} = 1.01 \times 10^{23}$ atoms P

Check: The units of the answer (atoms P) are correct. The magnitude of the answer is reasonable since there is less than the mass of 1 mol of P present.

(b) **Given:** 2.26 g Hg **Find:** atoms Hg
Conceptual Plan: g Hg → mol Hg → atoms Hg

$$\frac{1 \text{ mol Hg}}{200.59 \text{ g Hg}} \qquad \frac{6.022 \times 10^{23} \text{ atoms}}{\text{mol}}$$

Solution: $2.26 \cancel{\text{g Hg}} \times \frac{1 \cancel{\text{mol Hg}}}{200.59 \cancel{\text{g Hg}}} \times \frac{6.022 \times 10^{23} \text{ atoms Hg}}{1 \cancel{\text{mol Hg}}} = 6.78 \times 10^{21}$ atoms Hg

Check: The units of the answer (atoms Hg) are correct. The magnitude of the answer is reasonable since there is much less than the mass of 1 mol of Hg present.

(c) **Given:** 1.87 g Bi **Find:** atoms Bi
Conceptual Plan: $\textbf{g Bi} \xrightarrow{\frac{1\ \text{mol Bi}}{208.98\ \text{g Bi}}} \textbf{mol Bi} \xrightarrow{\frac{6.022 \times 10^{23}\ \text{atoms}}{\text{mol}}} \textbf{atoms Bi}$

Solution: $1.87\ \cancel{\text{g Bi}} \times \dfrac{1\ \cancel{\text{mol Bi}}}{208.98\ \cancel{\text{g Bi}}} \times \dfrac{6.022 \times 10^{23}\ \text{atoms Bi}}{1\ \cancel{\text{mol Bi}}} = 5.39 \times 10^{21}\ \text{atoms Bi}$

Check: The units of the answer (atoms Bi) are correct. The magnitude of the answer is reasonable since there is less than the mass of 1 mol of Bi present.

(d) **Given:** 0.082 g Sr **Find:** atoms Sr
Conceptual Plan: $\textbf{g Sr} \xrightarrow{\frac{1\ \text{mol Sr}}{87.62\ \text{g Sr}}} \textbf{mol Sr} \xrightarrow{\frac{6.022 \times 10^{23}\ \text{atoms}}{\text{mol}}} \textbf{atoms Sr}$

Solution: $0.082\ \cancel{\text{g Sr}} \times \dfrac{1\ \cancel{\text{mol Sr}}}{87.62\ \cancel{\text{g Sr}}} \times \dfrac{6.022 \times 10^{23}\ \text{atoms Sr}}{1\ \cancel{\text{mol Sr}}} = 5.6 \times 10^{20}\ \text{atoms Sr}$

Check: The units of the answer (atoms Sr) are correct. The magnitude of the answer is reasonable since there is less than the mass of 1 mol of Sr present.

2.68 (a) **Given:** 14.955 g Cr **Find:** atoms Cr
Conceptual Plan: $\textbf{g Cr} \xrightarrow{\frac{1\ \text{mol Cr}}{52.00\ \text{g Cr}}} \textbf{mol Cr} \xrightarrow{\frac{6.022 \times 10^{23}\ \text{atoms}}{\text{mol}}} \textbf{atoms Cr}$

Solution: $14.955\ \cancel{\text{g Cr}} \times \dfrac{1\ \cancel{\text{mol Cr}}}{52.00\ \cancel{\text{g Cr}}} \times \dfrac{6.022 \times 10^{23}\ \text{atoms Cr}}{1\ \cancel{\text{mol Cr}}} = 1.732 \times 10^{23}\ \text{atoms Cr}$

Check: The units of the answer (atoms Cr) are correct. The magnitude of the answer is reasonable since there is less than the mass of 1 mol of Cr present.

(b) **Given:** 39.733 g S **Find:** atoms S
Conceptual Plan: $\textbf{g S} \xrightarrow{\frac{1\ \text{mol S}}{32.07\ \text{g S}}} \textbf{mol S} \xrightarrow{\frac{6.022 \times 10^{23}\ \text{atoms}}{\text{mol}}} \textbf{atoms S}$

Solution: $39.733\ \cancel{\text{g S}} \times \dfrac{1\ \cancel{\text{mol S}}}{32.07\ \cancel{\text{g S}}} \times \dfrac{6.022 \times 10^{23}\ \text{atoms S}}{1\ \cancel{\text{mol S}}} = 7.461 \times 10^{23}\ \text{atoms S}$

Check: The units of the answer (atoms S) are correct. The magnitude of the answer is reasonable since there is slightly more than the mass of 1 mol of S present.

(c) **Given:** 12.899 g Pt **Find:** atoms Pt
Conceptual Plan: $\textbf{g Pt} \xrightarrow{\frac{1\ \text{mol Pt}}{195.08\ \text{g Pt}}} \textbf{mol Pt} \xrightarrow{\frac{6.022 \times 10^{23}\ \text{atoms}}{\text{mol}}} \textbf{atoms Pt}$

Solution: $12.899\ \cancel{\text{g Pt}} \times \dfrac{1\ \cancel{\text{mol Pt}}}{195.08\ \cancel{\text{g Pt}}} \times \dfrac{6.0221 \times 10^{23}\ \text{atoms Pt}}{1\ \cancel{\text{mol Pt}}} = 3.892 \times 10^{22}\ \text{atoms Pt}$

Check: The units of the answer (atoms Pt) are correct. The magnitude of the answer is reasonable since there is less than the mass of 1 mol of Pt present.

(d) **Given:** 97.552 g Sn **Find:** atoms Sn
Conceptual Plan: $\textbf{g Sn} \xrightarrow{\frac{1\ \text{mol Sn}}{118.71\ \text{g Sn}}} \textbf{mol Sn} \xrightarrow{\frac{6.022 \times 10^{23}\ \text{atoms}}{\text{mol}}} \textbf{atoms Sn}$

Solution: $97.552\ \cancel{\text{g Sn}} \times \dfrac{1\ \cancel{\text{mol Sn}}}{118.71\ \cancel{\text{g Sn}}} \times \dfrac{6.022 \times 10^{23}\ \text{atoms Sn}}{1\ \cancel{\text{mol Sn}}} = 4.948 \times 10^{23}\ \text{atoms Sn}$

Check: The units of the answer (atoms Sn) are correct. The magnitude of the answer is reasonable since there is slightly less than the mass of 1 mol of Sn present.

2.69 (a) **Given:** 1.1×10^{23} gold atoms **Find:** grams Au
Conceptual Plan: $\textbf{atoms Au} \xrightarrow{\frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{atoms}}} \textbf{mol Au} \xrightarrow{\frac{196.97\ \text{g Au}}{1\ \text{mol Au}}} \textbf{g Au}$

Solution: $1.1 \times 10^{23}\ \cancel{\text{atoms Au}} \times \dfrac{1\ \cancel{\text{mol Au}}}{6.022 \times 10^{23}\ \cancel{\text{atoms Au}}} \times \dfrac{196.97\ \text{g Au}}{1\ \cancel{\text{mol Au}}} = 36\ \text{g Au}$

Check: The units of the answer (g Au) are correct. The magnitude of the answer is reasonable since there are fewer than Avogadro's number of atoms in the sample.

(b) **Given:** 2.82×10^{22} helium atoms **Find:** grams He

Conceptual Plan: atoms He → mol He → g He

$$\frac{1\text{ mol}}{6.022 \times 10^{23}\text{ atoms}} \qquad \frac{4.002\text{ g He}}{1\text{ mol He}}$$

Solution: $2.82 \times 10^{22}\ \cancel{\text{atoms He}} \times \dfrac{1\ \cancel{\text{mol He}}}{6.022 \times 10^{23}\ \cancel{\text{atoms He}}} \times \dfrac{4.002\text{ g He}}{1\ \cancel{\text{mol He}}} = 0.187\text{ g He}$

Check: The units of the answer (g He) are correct. The magnitude of the answer is reasonable since there are fewer than Avogadro's number of atoms in the sample.

(c) **Given:** 1.8×10^{23} lead atoms **Find:** grams Pb

Conceptual Plan: atoms Pb → mol Pb → g Pb

$$\frac{1\text{ mol}}{6.022 \times 10^{23}\text{ atoms}} \qquad \frac{207.2\text{ g Pb}}{1\text{ mol Pb}}$$

Solution: $1.8 \times 10^{23}\ \cancel{\text{atoms Pb}} \times \dfrac{1\ \cancel{\text{mol Pb}}}{6.022 \times 10^{23}\ \cancel{\text{atoms Pb}}} \times \dfrac{207.2\text{ g Pb}}{1\ \cancel{\text{mol Pb}}} = 62\text{ g Pb}$

Check: The units of the answer (g Pb) are correct. The magnitude of the answer is reasonable since there are fewer than Avogadro's number of atoms in the sample.

(d) **Given:** 7.9×10^{21} uranium atoms **Find:** grams U

Conceptual Plan: atoms U → mol U → g U

$$\frac{1\text{ mol}}{6.022 \times 10^{23}\text{ atoms}} \qquad \frac{238.029\text{ g U}}{1\text{ mol U}}$$

Solution: $7.9 \times 10^{21}\ \cancel{\text{atoms U}} \times \dfrac{1\ \cancel{\text{mol U}}}{6.022 \times 10^{23}\ \cancel{\text{atoms U}}} \times \dfrac{238.029\text{ g U}}{1\ \cancel{\text{mol U}}} = 3.1\text{ g U}$

Check: The units of the answer (g U) are correct. The magnitude of the answer is reasonable since there are fewer than Avogadro's number of atoms in the sample.

2.70 (a) **Given:** 7.55×10^{26} cadmium atoms **Find:** kg Cd

Conceptual Plan: atoms Cd → mol Cd → g Cd → kg Cd

$$\frac{1\text{ mol}}{6.022 \times 10^{23}\text{ atoms}} \qquad \frac{112.41\text{ g Cd}}{1\text{ mol Cd}} \qquad \frac{1\text{ kg}}{1000\text{ g}}$$

Solution:

$$7.55 \times 10^{26}\ \cancel{\text{atoms Cd}} \times \frac{1\ \cancel{\text{mol Cd}}}{6.022 \times 10^{23}\ \cancel{\text{atoms Cd}}} \times \frac{112.41\ \cancel{\text{g Cd}}}{1\ \cancel{\text{mol Cd}}} \times \frac{1\text{ kg Cd}}{1000\ \cancel{\text{g Cd}}} = 141\text{ kg Cd}$$

Check: The units of the answer (kg Cd) are correct. The magnitude of the answer is reasonable since there are many more than Avogadro's number of atoms in the sample.

(b) **Given:** 8.15×10^{27} nickel atoms **Find:** kg Ni

Conceptual Plan: atoms Ni → mol Ni → g Ni → kg Ni

$$\frac{1\text{ mol}}{6.022 \times 10^{23}\text{ atoms}} \qquad \frac{58.69\text{ g Ni}}{1\text{ mol Ni}} \qquad \frac{1\text{ kg}}{1000\text{ g}}$$

Solution:

$$8.15 \times 10^{27}\ \cancel{\text{atoms Ni}} \times \frac{1\ \cancel{\text{mol Ni}}}{6.022 \times 10^{23}\ \cancel{\text{atoms Ni}}} \times \frac{58.69\ \cancel{\text{g Ni}}}{1\ \cancel{\text{mol Ni}}} \times \frac{1\text{ kg Ni}}{1000\ \cancel{\text{g Ni}}} = 794\text{ kg Ni}$$

Check: The units of the answer (kg Ni) are correct. The magnitude of the answer is reasonable since there are many more than Avogadro's number of atoms in the sample.

(c) **Given:** 1.22×10^{27} manganese atoms **Find:** kg Mn

Conceptual Plan: atoms Mn → mol Mn → g Mn → kg Mn

$$\frac{1\text{ mol}}{6.022 \times 10^{23}\text{ atoms}} \qquad \frac{54.94\text{ g Mn}}{1\text{ mol Mn}} \qquad \frac{1\text{ kg}}{1000\text{ g}}$$

Solution:

$$1.22 \times 10^{27}\ \cancel{\text{atoms Mn}} \times \frac{1\ \cancel{\text{mol Mn}}}{6.022 \times 10^{23}\ \cancel{\text{atoms Mn}}} \times \frac{54.94\ \cancel{\text{g Mn}}}{1\ \cancel{\text{mol Mn}}} \times \frac{1\text{ kg Mn}}{1000\ \cancel{\text{g Mn}}} = 111\text{ kg Mn}$$

Check: The units of the answer (kg Mn) are correct. The magnitude of the answer is reasonable since there are many more than Avogadro's number of atoms in the sample.

(d) **Given:** 5.48 x 10^{29} lithium atoms **Find:** kg Li

Conceptual Plan: atoms Li → mol Li → g Li → kg Li

$$\frac{1\text{ mol}}{6.022 \times 10^{23}\text{ atoms}} \quad \frac{6.941\text{ g Li}}{1\text{ mol Li}} \quad \frac{1\text{ kg}}{1000\text{ g}}$$

Solution: $5.48 \times 10^{29}\ \cancel{\text{atoms Li}} \times \frac{1\ \cancel{\text{mol Li}}}{6.022 \times 10^{23}\ \cancel{\text{atoms Li}}} \times \frac{6.941\ \cancel{\text{g Li}}}{1\ \cancel{\text{mol Li}}} \times \frac{1\text{ kg Li}}{1000\ \cancel{\text{g Li}}} = 6.32 \times 10^{3}\text{ kg Li}$

Check: The units of the answer (kg Li) are correct. The magnitude of the answer is reasonable since there are many more than Avogadro's number of atoms in the sample.

2.71 **Given:** 52 mg diamond (carbon) **Find:** atoms C

Conceptual Plan: mg C → g C → mol C → atoms C

$$\frac{1\text{ g C}}{1000\text{ mg C}} \quad \frac{1\text{ mol C}}{12.011\text{ g C}} \quad \frac{6.022 \times 10^{23}\text{ atoms}}{\text{mol}}$$

Solution: $52\ \cancel{\text{mg C}} \times \frac{1\ \cancel{\text{g C}}}{1000\ \cancel{\text{mg C}}} \times \frac{1\ \cancel{\text{mol C}}}{12.011\ \cancel{\text{g C}}} \times \frac{6.022 \times 10^{23}\text{ atoms C}}{1\ \cancel{\text{mol C}}} = 2.6 \times 10^{21}\text{ atoms C}$

Check: The units of the answer (atoms C) are correct. The magnitude of the answer is reasonable since there is less than the mass of 1 mol of C present.

2.72 **Given:** 536 kg helium **Find:** atoms He

Conceptual Plan: kg He → g He → mol He → atoms He

$$\frac{1000\text{ g He}}{1\text{ kg He}} \quad \frac{1\text{ mol He}}{4.0026\text{ g He}} \quad \frac{6.022 \times 10^{23}\text{ atoms}}{\text{mol}}$$

Solution: $536\ \cancel{\text{kg He}} \times \frac{1000\ \cancel{\text{g He}}}{1\ \cancel{\text{kg He}}} \times \frac{1\ \cancel{\text{mol He}}}{4.0026\ \cancel{\text{g He}}} \times \frac{6.022 \times 10^{23}\text{ atoms He}}{1\ \cancel{\text{mol He}}} = 8.06 \times 10^{28}\text{ atoms He}$

Check: The units of the answer (atoms He) are correct. The magnitude of the answer is reasonable since there is much more than the mass of 1 mol of He present.

2.73 **Given:** 1 atom platinum **Find:** g Pt

Conceptual Plan: atoms Pt → mol Pt → g Pt

$$\frac{1\text{ mol}}{6.022 \times 10^{23}\text{ atoms}} \quad \frac{195.08\text{ g Pt}}{1\text{ mol Pt}}$$

Solution: $1\ \cancel{\text{atom Pt}} \times \frac{1\ \cancel{\text{mol Pt}}}{6.022 \times 10^{23}\ \cancel{\text{atoms Pt}}} \times \frac{195.08\text{ g Pt}}{1\ \cancel{\text{mol Pt}}} = 3.239 \times 10^{-22}\text{ g Pt}$

Check: The units of the answer (g Pt) are correct. The magnitude of the answer is reasonable since there is only 1 atom in the sample.

2.74 **Given:** 35 atoms xenon **Find:** g Xe

Conceptual Plan: atoms Xe → mol Xe → g Xe

$$\frac{1\text{ mol}}{6.022 \times 10^{23}\text{ atoms}} \quad \frac{131.29\text{ g Xe}}{1\text{ mol Xe}}$$

Solution: $35\ \cancel{\text{atom Xe}} \times \frac{1\ \cancel{\text{mol Xe}}}{6.022 \times 10^{23}\ \cancel{\text{atoms Xe}}} \times \frac{131.29\text{ g Xe}}{1\ \cancel{\text{mol Xe}}} = 7.631 \times 10^{-21}\text{ g Xe}$

Check: The units of the answer (g Xe) are correct. The magnitude of the answer is reasonable since there are only 35 atoms in the sample.

The Periodic Table and Atomic Mass

2.75
(a) K — Potassium is a metal
(b) Ba — Barium is a metal
(c) I — Iodine is a nonmetal
(d) O — Oxygen is a nonmetal
(e) Sb — Antimony is a metalloid

2.76	(a)	gold	Au is a metal
	(b)	fluorine	F is a nonmetal
	(c)	sodium	Na is a metal
	(d)	tin	Sn is a metal
	(e)	argon	Ar is a nonmetal
2.77	(a)	tellurium	Te is in group 16 and is a main group element
	(b)	potassium	K is in group 1 and is a main group element
	(c)	vanadium	V is in group 5 and is a transition element
	(d)	manganese	Mn is in group 7 and is a transition element
2.78	(a)	Cr	Chromium is in group 6 and is a transition element
	(b)	Br	Bromine is in group 17 and is a main group element
	(c)	Mo	Molybdenum is in group 6 and is a transition element
	(d)	Cs	Cesium is in group 1 and is a main group element
2.79	(a)	sodium	Na is in group 1 and is an alkali metal
	(b)	iodine	I is in group 17 and is a halogen
	(c)	calcium	Ca is in group 2 and is an alkaline earth metal
	(d)	barium	Ba is in group 2 and is an alkaline earth metal
	(e)	krypton	Kr is in group 18 and is a noble gas
2.80	(a)	F	Fluorine is in group 17 and is a halogen
	(b)	Sr	Strontium is in group 2 and is an alkaline earth metal
	(c)	K	Potassium is in group 1 and is an alkali metal
	(d)	Ne	Neon is in group 18 and is a noble gas
	(e)	At	Astatine is in group 17 and is a halogen

2.81 (a) N and Ni would not be similar. Nitrogen is a nonmetal, nickel is a metal.

(b) Mo and Sn would not be similar. Although both are metals, molybdenum is a transition metal and tin is a main group metal.

(c) Na and Mg would not be similar. Although both are main group metals, sodium is in group 1 and magnesium is in group 2.

(d) Cl and F would be most similar. Chlorine and fluorine are both in group 17. Elements in the same group have similar chemical properties.

(e) Si and P would not be similar. Silicon is a metalloid and phosphorus is a nonmetal.

2.82 (a) Nitrogen and oxygen would not be most similar. Although both are nonmetals, N is in group 15 and O is in group 16.

(b) Titanium and gallium would not be most similar. Although both are metals, Ti is a transition metal and Ga is a main group metal.

(c) Lithium and sodium would be most similar. Li and Na are both in group 1. Elements in the same group have similar chemical properties.

(d) Germanium and arsenic would not be most similar. Ge and As are both metalloids and would share some properties, but Ge is in group 14 and As is in group 15.

(e) Argon and bromine would not be most similar. Although both are nonmetals, Ar is in group 18 and Br is in group 17.

2.83 Main group metal atoms will lose electrons to form a cation with the same number of electrons as the nearest, previous noble gas.

Nonmetal atoms will gain electrons to form an anion with the same number of electrons as the nearest noble gas.

(a) O^{2-} O is a nonmetal and has 8 electrons. It will gain electrons to form an anion. The nearest noble gas is neon with 10 electrons, so O will gain 2 electrons.

(b) K^+ K is a main group metal and has 19 electrons. It will lose electrons to form a cation. The nearest noble gas is argon with 18 electrons, so K will lose 1 electron.

(c) Al^{3+} Al is a main group metal and has 13 electrons. It will lose electrons to form a cation. The nearest noble gas is neon with 10 electrons, so Al will lose 3 electrons.

(d) Rb^+ Rb is a main group metal and has 37 electrons. It will lose electrons to form a cation. The nearest noble gas is krypton with 36 electrons, so Rb will lose 1 electron.

2.84 Main group metal atoms will lose electrons to form a cation with the same number of electrons as the nearest, previous noble gas.

Nonmetal atoms will gain electrons to form an anion with the same number of electrons as the nearest noble gas.

(a) Mg^{2+} Mg is a main group metal and has 12 electrons. It will lose electrons to form a cation. The nearest noble gas is neon with 10 electrons, so Mg will lose 2 electrons.

(b) N^{3-} N is a nonmetal and has 7 electrons. It will gain electrons to form an anion. The nearest noble gas is neon with 10 electrons, so N will gain 3 electrons.

(c) F^- F is a nonmetal and has 9 electrons. It will gain electrons to form an anion. The nearest noble gas is neon with 10 electrons, so F will gain 1 electron.

(d) Na^+ Na is a main group metal and has 11 electrons. It will lose electrons to form a cation. The nearest noble gas is neon with 10 electrons, so Na will lose 1 electron.

Cumulative Problems

2.85 **Given:** 7.83 g HCN sample 1: 0.290 g H; 4.06 g N. 3.37 g HCN sample 2 **Find:** g C in sample 2

Conceptual Plan: g HCN sample 1 → g C in HCN sample 1 → ratio g C to g HCN → g C in HCN sample 2

$$\text{g HCN} - \text{g H} - \text{g N} \qquad \frac{\text{g C}}{\text{g HCN}} \qquad \text{g HCN} \times \frac{\text{g C}}{\text{g HCN}}$$

Solution: $7.83\text{ g HCN} - 0.290\text{ g H} - 4.06\text{ g N} = 3.48\text{ g C}$

$$3.37\,\cancel{\text{g HCN}} \times \frac{3.48\text{ g C}}{7.83\,\cancel{\text{g HCN}}} = 1.50\text{ g C}$$

Check: The units of the answer (g C) are correct. The magnitude of the answer is reasonable since the sample size is about half the original sample size, the g C are about half the original g C.

2.86 (a) **Given:** mass ratio S:O = 1.0:1.0 in SO_2 **Find:** mass ratio S:O in SO_3

Conceptual Plan: determine the ratio of O:O in SO_3 and SO_2 then determine g O per g S in SO_3

Solution: For a fixed amount of S, the ratio of O is $\frac{3\text{ O}}{2\text{ O}} = 1.5$. So, for 1 gram S, SO_3 would have 1.5 g O. The mass ratio of S:O = 1.0:1.5, that is 2.0:3.0, in SO_3.

Check: The answer is reasonable since the ratio is smaller than the ratio for SO_2 and SO_3 has to contain more O per gram of S.

(b) **Given:** mass ratio S:O = 1.0:1.0 in SO_2 **Find:** mass ratio S:O in S_2O
Conceptual Plan: determine the ratio of S:S in S_2O and SO_2 then determine g O per g S in S_2O
Solution: For a fixed amount of O, the ratio of S is $\frac{2\text{ S}}{0.5\text{ S}} = 4.0$. So, for 1 gram O, S_2O would have 4 gram S. The mass ratio of S:O = 4.0:1.0 in S_2O.
Check: The answer is reasonable since the ratio is larger than the ratio for SO_2, and S_2O has to contain more S per gram of O.

2.87 **Given:** in CO mass ratio O:C = 1.33:1; in compound X, mass ratio O:C = 2:1. **Find:** formula of X
Conceptual Plan: determine the mass ratio of O:O in the two compounds
Solution: For 1 gram of C $\dfrac{2\text{ g O in compound X}}{1.33\text{ g O in CO}} = 1.5$

So, the ratio of O to C in compound X has to be 1.5:1 and the formula is C_2O_3.
Check: The answer is reasonable since it fulfills the criteria of multiple proportions and the mass ratio of O:C is 2:1.

2.88 **Given:** mass ratio 1 atom N:1 atom^{12}C = 7:6; mass ratio 2 mol N:1 mol O in N_2O = 7:4
Find: mass of 1 mol O
Conceptual Plan: determine the mass ratio of O to ^{12}C from the mass ratio of N to ^{12}C and the mass ratio of N to O and then determine the mol ratio of ^{12}C to O; then use the mass of 1 mol ^{12}C to determine mass 1 mol O

$$\frac{12.00\text{ g }^{12}\text{C}}{1\text{ mol }^{12}\text{C}}$$

Solution: From the mass ratios, for every 7 grams N there are 6 grams ^{12}C, and for every 7 grams N there are 4 grams O. So, the mass ratio of O^{12}, C is 4:6.

$$\frac{1\,\cancel{\text{atom}\,^{12}\text{C}}}{1\,\cancel{\text{atom N}}} \times \frac{6.022 \times 10^{23}\,\cancel{\text{atom N}}}{1\,\cancel{\text{mol N}}} \times \frac{1\text{ mol }^{12}\text{C}}{6.022 \times 10^{23}\,\cancel{\text{atom}\,^{12}\text{C}}} \times \frac{2\,\cancel{\text{mol N}}}{1\text{ mol O}} = \frac{2\text{ mol }^{12}\text{C}}{1\text{ mol O}}$$

$$\frac{2\,\cancel{\text{mol}\,^{12}\text{C}}}{1\text{ mol O}} \times \frac{12.00\,\cancel{\text{g}\,^{12}\text{C}}}{1\,\cancel{\text{mol}\,^{12}\text{C}}} \times \frac{4\text{ g O}}{6\,\cancel{\text{g}\,^{12}\text{C}}} = 16.00\text{ g O/mol O}$$

Check: The units of the answer (g O/mol O) are correct. The magnitude of the answer is reasonable since it is close to the value on the periodic table.

2.89 **Given:** $^4\text{He}^{2+}$ = 4.00151 u **Find:** charge to mass ratio C/kg
Conceptual Plan: determine total charge on $^4\text{He}^{2+}$ and then u $^4\text{He}^{2+} \rightarrow$ g $^4\text{He}^{2+} \rightarrow$ kg $^4\text{He}^{2+}$

$$\frac{+1.60218 \times 10^{-19}\text{C}}{\text{proton}} \qquad \frac{1\text{ g}}{1.66054 \times 10^{-24}\text{ u}} \qquad \frac{1\text{ kg}}{1000\text{ g}}$$

Solution:
$$\frac{2\,\cancel{\text{protons}}}{1\text{ atom }^4\text{He}^{2+}} \times \frac{+1.60218 \times 10^{-19}\text{ C}}{\cancel{\text{proton}}} = \frac{3.20436 \times 10^{-19}\text{ C}}{\text{atom }^4\text{He}^{2+}}$$

$$\frac{4.00151\,\cancel{\text{u}}}{1\text{ atom }^4\text{He}^{2+}} \times \frac{1.66054 \times 10^{-24}\,\cancel{\text{g}}}{1\,\cancel{\text{u}}} \times \frac{1\text{ kg}}{1000\,\cancel{\text{g}}} = \frac{6.64466742 \times 10^{-27}\text{ kg}}{1\text{ atom }^4\text{He}^{2+}}$$

$$\frac{3.20436 \times 10^{-19}\text{ C}}{\cancel{\text{atom}\,^4\text{He}^{2+}}} \times \frac{1\,\cancel{\text{atom}\,^4\text{He}^{2+}}}{6.64466742 \times 10^{-27}\text{ kg}} = 4.82245 \times 10^{7}\text{ C kg}^{-1}$$

Check: The units of the answer (C kg^{-1}) are correct. The magnitude of the answer is reasonable when compared to the charge to mass ratio of the electron.

2.90 **Given:** 12.3849 g sample I; atomic mass I = 126.9045 u; 1.00070g ^{129}I; mass ^{129}I = 128.9050 u
Find: mass of contaminated sample
Conceptual Plan: total mass of sample $\rightarrow$ fraction I and ^{129}I in the sample $\rightarrow$ apparent "atomic mass"

$$\text{mass I} + \text{mass }^{129}\text{I} \qquad \frac{\text{g I}}{\text{g sample}}; \frac{\text{g }^{129}\text{I}}{\text{g sample}} \qquad \text{Atomic mass} = \sum_n (\text{fraction of isotope } n) \times (\text{mass of isotope } n)$$

Solution: 12.3849 g I + 1.00070 g ^{129}I = 13.3856 g sample

$$\frac{12.3849\text{ g}}{13.3856\text{ g}} = 0.925240557\text{ fraction I} \qquad \frac{1.00070\text{ g}}{13.3856\text{ g}} = 0.07475944\text{ fraction }^{129}\text{I}$$

$$\text{Atomic mass} = \sum_n (\text{fraction of isotope } n) \times (\text{mass of isotope } n)$$
$$= (0.925240\underline{5}57)(126.9045\text{ u}) + (0.074759\underline{4}4)(128.9050\text{ u})$$
$$= 127.055\text{ u}$$

Check: The units of the answer (u) are correct. The magnitude of the answer is reasonable because it is between 126.9045 and 128.9050 and only slightly higher than the naturally occurring value.

2.91 $^{236}_{90}Th$ A – Z = number of neutrons. 236 – 90 = 146 neutrons. So, any nucleus with 146 neutrons is an isotone of $^{236}_{90}Th$.

Some would be $^{237}_{91}Pa$; $^{238}_{92}U$; $^{239}_{93}Np$; $^{240}_{94}Pu$; $^{235}_{89}Ac$; $^{234}_{88}Ra$; $^{241}_{95}Am$; $^{244}_{98}Cf$; etc.

2.92

Symbol	Z	A	Number protons	Number electrons	Number neutrons	Charge
Si	14	**28**	**14**	14	14	**0**
S^{2-}	**16**	32	**16**	**18**	**16**	2 –
Cu^{2+}	**29**	**63**	**29**	**27**	34	2+
P	15	**31**	**15**	15	16	**0**

2.93

Symbol	Z	A	Number protons	Number electrons	Number neutrons	Charge
O^{2-}	8	**16**	**8**	**10**	8	2 –
Ca^{2+}	20	**40**	**20**	**18**	20	**2+**
Mg^{2+}	**12**	25	**12**	**10**	13	**2+**
N^{3-}	**7**	14	**7**	10	**7**	**3 –**

2.94 **Given:** r (neutron) = 1.0×10^{-13} cm; r (star piece) = 0.10 mm **Find:** density of neutron, mass (kg) of star piece
Conceptual Plan: **r (neutron) $\xrightarrow{V = \frac{4}{3}\pi r^3}$ vol (neutron) $\xrightarrow{d = \frac{m}{v}}$ density (neutron) and then r (star piece) $\xrightarrow{V = \frac{4}{3}\pi r^3}$ vol (star piece) $\xrightarrow{m = dv}$ mass (star piece)**

Solution:

For the neutron:

$$\text{Vol(neutron)} = \frac{4}{3}\pi(1.0 \times 10^{-13}\text{cm})^3 = 4.19 \times 10^{-39}\text{cm}^3 \qquad d = \frac{1.00727\,\cancel{\text{u}}}{4.19 \times 10^{-39}\text{cm}^3} \times \frac{1.661 \times 10^{-24}\text{g}}{\cancel{\text{u}}}$$
$$= 3.99 \times 10^{14}\text{g cm}^{-3}$$

For the star piece:

$$\text{Vol(star piece)} = \frac{4}{3}\pi(0.10\,\cancel{\text{mm}})^3 \frac{(1\,\text{cm})^3}{(10\,\cancel{\text{mm}})^3} = 4.19 \times 10^{-4}\text{cm}^3 \qquad m = 4.19 \times 10^{-4}\cancel{\text{cm}^3} \times \frac{3.99 \times 10^{14}\cancel{\text{g}}}{\cancel{\text{cm}^3}} \times \frac{1\,\text{kg}}{1000\,\cancel{\text{g}}}$$
$$= 1.7 \times 10^{8}\text{kg}$$

Check: The units of the answer (kg) are correct. The magnitude of the answer shows the great mass of the neutron star.

2.95 **Given:** r(nucleus) = 2.7 fm; r(atom) = 70 pm (assume two significant figures)
Find: vol(nucleus); vol(atom); % vol(nucleus)

Conceptual Plan:
r(nucleus)(fm) → r(nucleus)(pm) → vol(nucleus) and then r(atom) → vol(atom) and then % vol

$\frac{10^{-15}\,\text{m}}{1\,\text{fm}}$ $\frac{1\,\text{pm}}{10^{-12}\,\text{m}}$ $V = \frac{4}{3}\pi r^3$ $V = \frac{4}{3}\pi r^3$ $\frac{\text{vol(nucleus)}}{\text{vol(atom)}} \times 100\%$

Solution:

$$2.7\,\cancel{\text{fm}} \times \frac{10^{-15}\,\cancel{\text{m}}}{\cancel{\text{fm}}} \times \frac{1\text{ pm}}{10^{-12}\,\cancel{\text{m}}} = 2.7 \times 10^{-3}\text{pm} \qquad V_{\text{nucleus}} = \frac{4}{3}\pi\,(2.7 \times 10^{-3}\text{pm})^3 = 8.2 \times 10^{-8}\text{ pm}^3$$

$$V_{\text{atom}} = \frac{4}{3}\pi\,(70\text{ pm})^3 = 1.4 \times 10^6\text{ pm}^3 \qquad \frac{8.2 \times 10^{-8}\,\cancel{\text{pm}^3}}{1.4 \times 10^6\,\cancel{\text{pm}^3}} \times 100\% = 5.9 \times 10^{-12}\%$$

Check: The units of the answer (% vol) are correct. The magnitude of the answer is reasonable because the nucleus only occupies a very small % of the vol of the atom.

2.96 **Given:** mass of ^{14}N = 14.003074 u, mass of ^{15}N = 15.000109 u, atomic mass of N = 14.0067 u
Find: abundance fractions of ^{14}N and ^{15}N
Conceptual Plan: x = fraction of ^{14}N, y = fraction of ^{15}N. mass of ^{14}N, mass of ^{15}N, atomic mass of N → x, y

Solution: Atomic mass = 14.0067 u = x(14.003074 u) + y(15.000109 u)

$y = 1 - x$

Therefore,

$14.0067\text{ u} = x(14.003074\text{ u}) + (1 - x)(15.000109\text{ u})$

$14.0067 = 14.003074x - 15.000109x + 15.000109$

$0.993\underline{4}09 = 0.997035x$

$x = 0.993\underline{4}09/0.997035 = 0.9964;\ \%^{14}N = 0.9964 \times 100 = 99.64\%$

$y = 1 - 0.9964 = 0.0036;\ \%^{15}N = 0.0036 \times 100 = 0.36\%$

Check: The answer makes sense, because the average atomic mass of nitrogen is very close to that of ^{14}N, so one would expect that this isotope must have the overwhelming abundance.

2.97 **Given:** mass of ^{12}C = 12 exactly by definition **Find:** atomic mass of O if atomic mass of C was used instead of mass of ^{12}C
Other: atomic mass of C = 12.0107 u, atomic mass of O = 15.9994 u
Conceptual Plan: atomic mass of O/atomic mass of C → q. 12.000, q → atomic mass of O (using the new scale)
Solution: Mass of an isotope is determined in relationship to mass of ^{12}C. The mass ratio between the isotope of interest and ^{12}C is multiplied by 12 to determine the atomic mass of that isotope. For the sake of this problem, we are assuming that oxygen is a single isotope and its mass = 15.9994 u. Since we are using atomic mass of C as the benchmark:

$$\frac{\text{mass O}}{\text{mass C}} = \frac{15.9994\text{ u}}{12.0107\text{ u}} = 1.3320\underline{9}5548$$

Now, as the problem requested, we will arbitrarily assign the mass of 12.0107 as 12.0000. But this does not change the true ratio between atomic mass of C and that of O. But as before, mass of O is determined relative to mass of C. Therefore,

$\text{mass O} = \text{mass C} \times q = 12.000\text{ u} \times 1.3320\underline{9}5548 = 15.985\text{ u}$

Check: Since the atomic mass of carbon, 12.0107, is closer to that of O, their ratio would be smaller than between ^{12}C and O. Therefore, the calculated mass would be smaller than the actual atomic mass of O.

2.98 **Given:** Ti cube: d = 4.50 g cm^{-3}; e = 2.78 cm **Find:** number Ti atoms
Conceptual Plan: → vol cube → g Ti → mol Ti → atoms Ti

$V = e^3$ $\frac{4.50\text{ g}}{\text{cm}^3}$ $\frac{1\text{ mol Ti}}{47.87\text{ g}}$ $\frac{6.022 \times 10^{23}\text{ atoms}}{\text{mol}}$

Solution: $(2.78\ \cancel{cm})^3 \times \frac{4.50\ \cancel{g}}{\cancel{cm^3}} \times \frac{1\ \cancel{mol\ Ti}}{47.87\ \cancel{g}} \times \frac{6.022 \times 10^{23}\ \text{atoms Ti}}{1\ \cancel{mol\ Ti}} = 1.22 \times 10^{24}\ \text{atoms Ti}$

Check: The units of the answer (atoms Ti) are correct. The magnitude of the answer is reasonable because there is about 2 mol of Ti in the cube.

2.99 **Given:** Cu sphere: r = 0.935 cm; d = 8.94 g cm^{-3} **Find:** number of Cu atoms
Conceptual Plan: r in inch → r in cm → vol sphere → g Cu → mol Cu → atoms Cu

$V = \frac{4}{3}\pi r^3 \qquad \frac{8.96\ g}{cm^3} \qquad \frac{1\ mol\ Cu}{63.546\ g} \qquad \frac{6.022 \times 10^{23}\ atoms}{mol}$

Solution:

$$\frac{4}{3}\pi(0.935)^3 \times \frac{8.94\ \cancel{g}}{\cancel{cm^3}} \times \frac{1\ \cancel{mol\ Cu}}{63.546\ \cancel{g}} \times \frac{6.022 \times 10^{23}\ \text{atoms Cu}}{1\ \cancel{mol\ Cu}} = 2.90 \times 10^{23}\ \text{atoms Cu}$$

Check: The units of the answer (atoms Cu) are correct. The magnitude of the answer is reasonable because there are about 8 mol Cu present.

2.100 **Given:** B-10 = 10.01294 u; B-11 = 11.00931 u; B = 10.81 u **Find:** % abundance B-10 and B-11
Conceptual Plan: Let x = fraction B-10 then 1 – x = fraction B-11 → abundances

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

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

$10.81 = (x)(10.01294\ u) + (1 - x)(11.00931\ u)$

$0.1\underline{9}931 = 0.99637\,x$

$x = 0.2\underline{0}0 \qquad 1 - x = 0.8\underline{0}0$

B-10 = $0.2\underline{0}0 \times 100$ = 20. % and B-11 = $0.8\underline{0}0 \times 100$ = 80. %

Check: The units of the answer (%, which gives the relative abundance of each isotope) are correct. The relative abundances are reasonable because B has an atomic mass closer to the mass of B-11 than to B-10.

2.101 **Given:** Li-6 = 6.01512 u; Li-7 = 7.01601 u; B = 6.941 u **Find:** % abundance Li-6 and Li-7
Conceptual Plan: Let x = fraction Li-6 then 1 – x = fraction Li-7 → abundances

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

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

$6.941 = (x)(6.01512\ u) + (1 - x)(7.01601\ u)$

$0.07501 = 1.00089\,x$

$x = 0.07494 \qquad 1 - x = 0.92506$

Li-6 = 0.07494 x 100 = 7.494% and Li-7 = 0.92506 x 100 = 92.506%

Check: The units of the answer (%, which gives the relative abundance of each isotope) are correct. The relative abundances are reasonable because Li has an atomic mass closer to the mass of Li-7 than to Li-6.

2.102 **Given:** Brass: 37.0% Zn, d = 8.48 g cm^{-3}, volume = 112.5 cm^3 **Find:** atoms of Zn and Cu
Conceptual Plan: Volume sample → g sample → g Zn → mole Zn → atoms Zn

$\frac{8.48\ g}{cm^3} \qquad \frac{37.0\ g\ Zn}{100.0\ g\ sample} \qquad \frac{65.41\ g\ Zn}{mol\ Zn} \qquad \frac{6.022 \times 10^{23}\ atoms}{mol}$

→ g Cu → moles Cu → atoms Cu

$g\ sample - g\ Zn \qquad \frac{63.55\ g\ Cu}{mol\ Cu} \qquad \frac{6.022 \times 10^{23}\ atoms}{mol}$

Solution: $112.5\ \cancel{cm^3} \times \frac{8.48\ g}{\cancel{cm^3}} = 95\underline{4}.0\ g\ sample \qquad 95\underline{4}.0\ \cancel{g\ sample} \times \frac{37.0\ g\ Zn}{100.0\ \cancel{g\ sample}} = 35\underline{2}.98\ g\ Zn$

$$352.98\,\cancel{\text{g Zn}} \times \frac{1\,\cancel{\text{mol Zn}}}{65.41\,\cancel{\text{g Zn}}} \times \frac{6.022 \times 10^{23}\text{atoms Zn}}{\cancel{\text{mol Zn}}} = 3.2497 \times 10^{24}\text{atoms Zn} = 3.25 \times 10^{24}\text{atoms Zn}$$

$$954.0\,\text{g sample} - 352.98\text{ g Zn} = 601.02\text{ g Cu}$$

$$601.01\,\cancel{\text{g Cu}} \times \frac{1\,\cancel{\text{mol Cu}}}{63.55\,\cancel{\text{g Cu}}} \times \frac{6.022 \times 10^{23}\text{atoms Cu}}{\cancel{\text{mol Cu}}} = 5.6952 \times 10^{24}\text{atoms Cu} = 5.70 \times 10^{24}\text{atoms Cu}$$

Check: The units of the answer (atoms of Zn and atoms of Cu) are correct. The magnitude is reasonable since there is more than 1 mole of each element in the sample.

2.103 **Given:** alloy of Au and Pd = 67.2 g; 2.49 x 10^{23} atoms **Find:** % composition by mass

Conceptual Plan: atoms Au and Pd → mol Au and Pd → g Au and Pd → g Au

$$\frac{1\text{ mol}}{6.022 \times 10^{23}\text{atoms}} \qquad \frac{196.97\text{g Au}}{1\text{ mol Au}};\ \frac{106.42\text{ g Pd}}{1\text{ mol Pd}}$$

Solution: Let X = atoms Au and Y = atoms Pd, develop expressions that will permit atoms to be related to moles and then to grams.

$$(X\,\cancel{\text{atoms Au}})\left(\frac{1\text{ mol Au}}{6.022 \times 10^{23}\cancel{\text{atoms Au}}}\right) = \frac{X}{6.022 \times 10^{23}}\text{ mol Au}$$

$$(Y\,\cancel{\text{atoms Pd}})\left(\frac{1\text{ mol Pd}}{6.022 \times 10^{23}\cancel{\text{atoms Pd}}}\right) = \frac{Y}{6.022 \times 10^{23}}\text{ mol Pd}$$

X + Y = 2.49 x 10^{23} atoms; Y = 2.49 x 10^{23} – X

$$\left(\frac{X}{6.022 \times 10^{23}}\cancel{\text{mol Au}}\right)\left(\frac{196.97\text{ g Au}}{\cancel{\text{mol Au}}}\right) = \frac{196.97X}{6.022 \times 10^{23}}\text{ g Au}$$

$$\left(\frac{2.49 \times 10^{23} - X}{6.022 \times 10^{23}}\cancel{\text{mol Pd}}\right)\left(\frac{106.42\text{ g Pd}}{\cancel{\text{mol Pd}}}\right) = \frac{106.42(2.49 \times 10^{23} - X)}{6.022 \times 10^{23}}\text{ g Pd}$$

g Au + g Pd = 67.2 g total

$$\frac{196.97X}{6.022 \times 10^{23}}\text{ g Au} + \frac{106.42(2.49 \times 10^{23} - X)}{6.022 \times 10^{23}}\text{ g Pd} = 67.2\text{ g}$$

X = 1.5426 x 10^{23} atoms Au

$$(1.54 \times 10^{23}\cancel{\text{atoms Au}})\left(\frac{1\,\cancel{\text{mol Au}}}{6.022 \times 10^{23}\cancel{\text{atoms Au}}}\right)\left(\frac{196.97\text{ g Au}}{\cancel{\text{mol Au}}}\right) = 50.37\text{ g Au}$$

$$\left(\frac{50.37\text{ g Au}}{67.2\text{ g sample}}\right) \times 100 = 74.95\%\text{ Au} = 75.0\%\text{ Au}$$

$$\%\text{ Pd} = 100.0\% - 75.0\%\text{ Au} = 25.0\%\text{ Pd}$$

Check: Units of the answer (% composition) are correct.

2.104 **Given:** Cl-35, mass = 34.9688 u, 75.76%; Cl-37, mass = 36.9659 u, 24.24%; O-16, mass = 15.9949 u, 99.57%; O-17, mass = 16.9991 u, 0.038%; O-18, mass = 17.9991, 0.205%

Find: number of different masses of Cl_2O, the mass of the three most abundant

Conceptual Plan: Determine the different combinations of Cl and O. Use the % abundance to determine the most abundant. Determine the mass of the molecule.

$$\text{Mass} = \sum \text{mass of each isotope}$$

Solution: Possible combinations:

$^{35}Cl^{35}Cl^{16}O$	$^{35}Cl^{35}Cl^{17}O$	$^{35}Cl^{35}Cl^{18}O$
$^{35}Cl^{37}Cl^{16}O$	$^{35}Cl^{37}Cl^{17}O$	$^{35}Cl^{37}Cl^{18}O$
$^{37}Cl^{37}Cl^{16}O$	$^{37}Cl^{37}Cl^{17}O$	$^{37}Cl^{37}Cl^{18}O$

So, there are nine possible combinations and nine different masses of Cl_2O.

O-17 and O-18 are both less than 1% naturally occurring, so molecules containing these isotopes will not be very abundant, therefore, the three most abundant molecules will be the ones that contain O-16; $^{35}Cl^{35}Cl^{16}O$, $^{35}Cl^{37}Cl^{16}O$, $^{37}Cl^{37}Cl^{16}O$.

Mass $^{35}Cl^{35}Cl^{16}O$ = 34.9688 u + 34.9688 u + 15.9949 u = 85.9325 u

Mass $^{35}Cl^{37}Cl^{16}O$ = 34.9688 u + 36.9659 u + 15.9949 u = 87.9296 u

Mass $^{37}Cl^{37}Cl^{16}O$ = 36.9659 u + 36.9659 u + 15.9949 u = 89.9267 u

Check: The units of the answer (u) are correct.

2.105 **Given:** Ag-107, 51.839%, Ag-109, $\frac{\text{mass Ag-109}}{\text{mass Ag-107}} = 1.0187$ **Find:** mass Ag-107

Conceptual Plan: % abundance Ag-107 → % abundance Ag-109 → fraction → mass Ag-107

$100\% - (\%\text{Ag-107})$ $\quad \frac{\%\text{abundance}}{100}$

Atomic mass $= \sum_n$(fraction of isotope n) x (mass of isotope n)

Solution: 100.00% – 51.839% = 48.161% Ag – 109

$\text{Fraction Ag-107} = \frac{51.839}{100.00} = 0.51839 \qquad \text{Fraction Ag-109} = \frac{48.161}{100.00} = 0.48161$

Let X be the mass of Ag-107 then mass Ag-109 = 1.0187X

Atomic mass $= \sum_n$(fraction of isotope n) x (mass of isotope n)

$107.87\text{ u} = 0.51839(\text{X u}) + 0.48161(1.0187\text{X u})$

$\text{X} = 106.9\underline{0}7\text{ u} = 106.91\text{ u mass Ag-107}$

Check: The units of the answer (u) are correct. The answer is reasonable since it is close to the atomic mass number of Ag-107.

Challenge Problems

2.106 (a) **Given:** 36 Sk-296; 2 Sk-297; 12 Sk-298 **Find:** % abundance of each

Conceptual Plan: total atoms → fraction of each isotope → % abundance

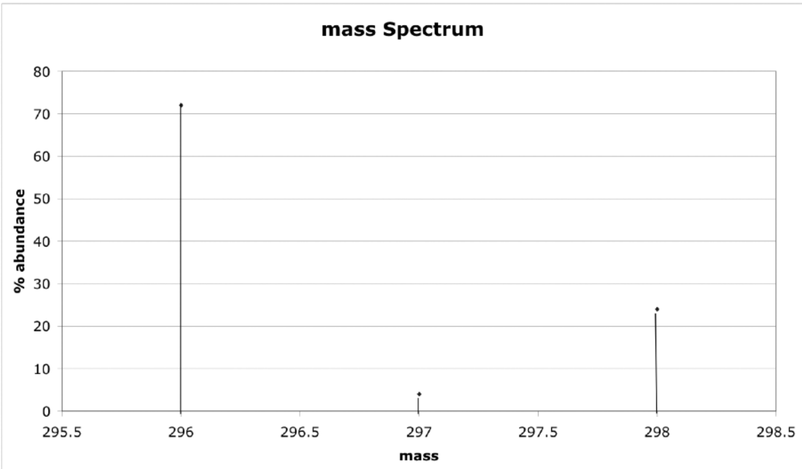
Sum of atoms $\quad \frac{\text{number of each isotope}}{\text{total atoms}} \quad$ fraction x 100%

Solution: Total atoms = 36 + 2 + 12 = 50

$\frac{36}{50} \text{x} 100\% = 72\%\text{Sk-296}, \quad \frac{2}{50} \text{x} 100\% = 4\%\text{Sk-297}, \quad \frac{12}{50} \text{x} 100\% = 24\%\text{Sk-298}$

Check: The units of the answers (% abundance) are correct. The values of the answers are reasonable since they add up to 100%

(b)



(c) **Given:** Sk-296 m = 24.6630 x mass ^{12}C, 72.55%: Sk-297 m = 24.7490 x mass ^{12}C, 3.922%: Sk-298 m = 24.8312 x mass ^{12}C; 23.53%.

Find: atomic mass Sk

Conceptual Plan: mass of isotope relative to ^{12}C → mass of isotope and then % abundance →

(Mass relative to ^{12}C)(12.00 u) $\quad \frac{\%\text{abundance}}{100}$

fraction abundance then determine atomic mass

Atomic mass $= \sum_n$(fraction of isotope n) x (mass of isotope n)

Solution:

Sk-296 = 24.6630 x 12.00 u = 295.956 u; Sk-297 = 24.7490 x 12.00 u = 296.988 u;

Sk-298 = 24.8312 x 12.00 u = 297.974 u

$$\text{fraction Sk-296} = \frac{72}{100} = 0.72 \quad \text{fraction Sk-297} = \frac{4}{100} = 0.04 \quad \text{fraction Sk-298} = \frac{24}{100} = 0.24$$

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

$$= (0.72)(295.956 \text{ u}) + (0.04)(296.988 \text{ u}) + (0.24)(297.974 \text{ u})$$

$$= 296.482 \text{ u}$$

Check: The units of the answer (u) are correct. The magnitude of the answer is reasonable because it lies between 295.956 and 297.974 and is closer to 296, which has the highest abundance.

2.107 **Given:** $\frac{\text{mass 2 O}}{\text{mass 1 N}} = \frac{2.29}{1.00}; \frac{\text{mass 3 F}}{\text{mass 1 N}} = \frac{4.07}{1.00}$ **Find:** $\frac{\text{mass O}}{\text{mass 2 F}}$

Conceptual Plan: mass O/N and mass F/N $\longrightarrow$ mass O/F $\rightarrow$ mass O/2F

$$\frac{\text{mass 2 O}}{\text{mass 1 N}} \quad \frac{\text{mass 3 F}}{\text{mass 1 N}} \quad \frac{\text{mass 2 O}}{\text{mass 3 F}}$$

Solution: $\frac{\text{mass 2 O}}{\text{mass 1 N}} = \frac{2.29}{1.00}; \frac{\text{mass 3 F}}{\text{mass 1 N}} = \frac{4.07}{1.00} \quad \left(\frac{2.29 \text{ mass } \cancel{2\text{ O}}}{4.07 \text{ mass } \cancel{3\text{ F}}}\right)\left(\frac{1\text{O}}{\cancel{2\text{O}}}\right)\left(\frac{\cancel{3\text{ F}}}{2\text{ F}}\right) = \frac{0.422 \text{ mass O}}{\text{mass 2 F}}$

Check: Mass ratio of O to F is reasonable since the mass of O is slightly less than the mass of fluorine.

2.108 **Given:** Sample = 1.5886 g, ^{59}Co = 58.9332 u, ^{60}Co = 59.9338 u, apparent mass = 58.9901 u
Find: mass of ^{60}Co in sample
Conceptual Plan: apparent mass $\rightarrow$ fraction ^{60}Co $\rightarrow$ mass ^{60}Co

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

Solution: Let X = fraction of ^{60}Co, so: 1.00 – X = fraction ^{59}Co

$$58.9901 \text{ u} = (1.00 - \text{X})(58.9332 \text{ u}) + (\text{X})(59.9338 \text{ u})$$

$$\text{X} = 0.056\underline{8}6$$

$$1.5886 \text{ g sample} \times 0.056\underline{8}6 = 0.090\underline{3}37 \text{ g } ^{60}\text{Co} = 0.0903 \text{ g } ^{60}\text{Co}$$

Check: The units of the answer (g ^{60}Co) are correct. The magnitude of the answer is reasonable since the apparent mass is very close to the mass ^{59}Co.

2.109 **Given:** 7.36 g Cu, 0.51 g Zn **Find:** atomic mass of sample
Conceptual Plan: fraction Cu and Zn $\rightarrow$ atomic mass

$$\text{Atomic mass} = \sum_n (\text{fraction of atom } n) \times (\text{mass of atom } n)$$

Solution: 7.36 g Cu + 0.51 g Zn = 7.87 g sample

$$\left(\frac{7.36 \text{ g Cu}}{7.87 \text{ g sample}}\right)\left(\frac{63.55 \text{ g Cu}}{\text{mol Cu}}\right) + \left(\frac{0.51 \text{ g Zn}}{7.87 \text{ g sample}}\right)\left(\frac{65.41 \text{ g Zn}}{\text{mol Zn}}\right) = 63.67 \text{ g mol}^{-1}$$

Check: Units of the answer (g mol^{-1}) are correct. The magnitude of the answer is reasonable since it is between the mass of Cu (63.55 g mol^{-1}) and Zn (65.41 g mol^{-1}) and is closer to the mass of Cu.

2.110 **Given:** $N_2O_3 = \frac{\text{mass O}}{\text{mass N}} = \frac{12}{7}$; sample X = $\frac{\text{mass O}}{\text{mass N}} = \frac{16}{7}$ **Find:** Formula of X, next in series
Conceptual Plan: ratio O/N for N_2O_3 $\rightarrow$ ratio O/N for X $\rightarrow$ ratio if O/O

Solution: $\frac{\text{mass O}}{\text{mass N}} = \frac{12}{7} = \frac{3\text{ O mass O}}{2\text{ N mass N}} = \frac{16}{7} = \frac{\text{X O mass O}}{2\text{ N mass O}} = \frac{16}{12} = \frac{\text{X O}}{3\text{ O}} \text{X} = 4$

Therefore, formula is N_2O_4

The next member of the series would be N_2O_5

$$\frac{\text{mass O}}{\text{mass O}} = \frac{5\text{ O}}{3\text{ O}} = \frac{\text{Y}}{12} \text{Y} = 20$$

$$\text{So, } \frac{\text{mass O}}{\text{mass N}} = \frac{20}{7}$$

2.111 **Given:** Mg = 24.312 u, ^{24}Mg = 23.98504, 78.99%, ^{26}Mg = 25.98259 u, $\dfrac{\text{abundance } ^{25}\text{Mg}}{\text{abundance } ^{26}\text{Mg}} = \dfrac{0.9083}{1}$

Find: mass ^{25}Mg

Conceptual Plan: abundance of ^{24}Mg and ratio ^{25}Mg /^{26}Mg $\longrightarrow$ abundance ^{25}Mg and ^{26}Mg $\rightarrow$ mass ^{25}Mg

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

Solution: 100.00% – % abundance ^{24} = % abundance ^{25}Mg and ^{26}Mg

100.00% – 78 99% = 21.01% ^{25}Mg and ^{26}Mg

fraction ^{25}Mg and ^{26}Mg = $\dfrac{21.01}{100.0} = 0.2101$

$\dfrac{\text{abundance } ^{25}\text{Mg}}{\text{abundance } ^{26}\text{Mg}} = \dfrac{0.9083}{1}$

Let X = fraction ^{26}Mg, 0.9083X = fraction ^{25}Mg

fraction ^{25}Mg and ^{26}Mg = X + 0.9083X = 0.2101

X = ^{26}Mg = 0.1101, 0.9083X = ^{25}Mg = 0.1000

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

24.312 = (0.7899)(23.98504 u) + (0.1000)(mass ^{25}Mg) + (0.1101)(25.98259 u)

mass ^{25}Mg = 25.0$\underline{5}$6 u = 25.06 u

Check: The units of the answer (u) are correct. The magnitude of the answer is reasonable since it is between the masses of ^{24}Mg and ^{26}Mg.

Conceptual Problems

2.112 (a) This is the law of definite proportions: All samples of a given compound, regardless of their source or how they were prepared, have the same proportions of their constituent elements.

(b) This is the law of conservation of mass: In a chemical reaction, matter is neither created nor destroyed.

(c) This is the law of multiple proportions: When two elements form two different compounds, the masses of element B that combine with 1 g of element A can be expressed as a ratio of small whole numbers. In this example the ratio of O from hydrogen peroxide to O from water = 16:8 $\rightarrow$ 2:1, a small whole number ratio.

2.113 If the u and mole were not based on the same isotope, the numerical values obtained for an atom of material and a mole of material would not be the same. If, for example, the mole was based on the number of particles in C – 12 but the u was changed to a fraction of the mass of an atom of Ne – 20 the number of particles and the number of u that make up one mole of material would no longer be the same. We would no longer have the relationship where the mass of an atom in u is numerically equal to the mass of a mole of those atoms in grams.

2.114 **Given:** a. Cr: 55.0 g, atomic mass = 52 g mol^{-1}; b. Ti: 45.0 g, atomic mass = 48 g mol^{-1}; and c. Zn: 60.0 g, atomic mass = 65 g mol^{-1}

Find: which has the greatest mol, and which has the greatest mass

Cr would have the greatest mole amount of the elements. It is the only one whose mass is greater than the molar mass.

2.115 The different isotopes of the same element have the same number of protons and electrons, so the attractive forces between the nucleus and the electrons is constant and there is no difference in the radii of the isotopes. Ions, on the other hand, have a different number of electrons than the parent atom from which they are derived. Cations have fewer electrons than the parent atom. The attractive forces are greater because there is a larger positive charge in the nucleus than the negative charge in the electron cloud. So, cations are smaller than the parent atom from which they are derived. Anions have more electrons than the parent. The electron cloud has a greater negative charge than the nucleus, so the anions have larger radii than the parent.

Chapter 2. Atoms and Elements

Student Objectives

2.1 Imaging and Moving Individual Atoms

- Describe scanning tunnelling microscopy (STM) and how atoms are imaged on surfaces.
- Define **atom** and **element**.

2.2 Early Ideas About the Building Blocks of Matter

- Describe the earliest definitions of atoms and matter (Greeks).
- Know that greater emphasis on observation and the development of the scientific method led to the scientific revolution.

2.3 Modern Atomic Theory and the Laws That Led to It

- State and understand the law of conservation of mass (also from Section 1.2).
- State and understand the law of definite proportions.
- State and understand the law of multiple proportions.
- Know the four postulates of Dalton's atomic theory.

2.4 Atomic Structure

- Define **radioactivity**, **nucleus**, **proton**, and **neutron**.
- Understand Thomson's plum-pudding model and how Ernest Rutherford's gold-foil experiment refuted it by giving evidence for a nuclear structure of the atom.

2.5 Atomic Mass: The Average Mass of an Element's Atoms

- Calculate atomic mass from isotope masses and natural abundances.
- Define **mass spectrometry** and understand how it can be used to measure mass and relative abundance.

2.6 Molar Mass: Counting Atoms by Weighing Them

- Understand the relationship between mass and count of objects such as atoms.
- Define **mole** and **Avogadro's number**.
- Calculate and interconvert between number of moles and atoms.
- Calculate and interconvert between number of moles and mass.

2.7 The Periodic Table of the Elements

- Define the **periodic law**.
- Know that elements with similar properties are placed into columns (called groups) in the periodic table.
- Define and distinguish between metals, nonmetals, and metalloids.
- Identify main-group and transition elements on the periodic table.
- Know the general properties of elements in some specific groups: noble gases, alkali metals, alkaline earth metals, and halogens.
- Know and understand the rationale for elements that form ions with predictable charges.

Chapter 2. Atoms and Elements

Section Summaries

Lecture Outline

- Terms, Concepts, Relationships, Skills
- Figures, Tables, and Solved Examples

Teaching Tips

- Suggestions and Examples
- Misconceptions and Pitfalls

Chapter 2. Atoms and Elements

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples
2.1 Imaging and Moving Individual Atoms • Description of scanning tunnelling microscopy (STM) • Introduction to macroscopic and microscopic perspectives • Definitions of atom and element	• Intro figure: tip of an STM moving across a surface • Figure 2.1: Scanning Tunnelling Microscopy • Figure 2.2: Imaging Atoms
2.2 Early Ideas About the Building Blocks of Matter • History of chemistry from antiquity (~450 BC) • Scientific revolution (1400s–1600s)	

2.3 Modern Atomic Theory and the Laws That Led to It
• Law of conservation of mass
  o Matter is neither created nor destroyed.
  o Atoms at the start of a reaction may recombine to form different compounds, but all atoms are accounted for at the end.
  o Mass of reactants = mass of products.
• Law of definite proportions
  o Different samples of the same compound have the same proportions of constituent elements independent of sample source or size.
• Law of multiple proportions
• John Dalton's atomic theory
	• Unnumbered figure: models and photos of Na and Cl_2 forming NaCl
	• Example 2.1: Law of Definite Proportions
	• Unnumbered figure: space filling models of CO and CO_2 illustrating the law of multiple proportions

Chapter 2. Atoms and Elements

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.1 Imaging and Moving Individual Atoms • Other STM images can be found readily on the Internet. • It is useful to reiterate the analogies about size; the one used in the chapter compares an atom to a grain of sand and a grain of sand to a large mountain range.	• STM is not actually showing images of atoms like one might imagine seeing with a light microscope. • Atoms are not coloured spheres; the images use colour to distinguish different atoms.

2.2 Early Ideas About the Building Blocks of Matter • The view of matter as made up of small, indestructible particles was ignored because more popular philosophers like Aristotle and Socrates had different views. • Leucippus and Democritus may have been proven correct, but they had no more evidence for their ideas than Aristotle did. • Observations and data led scientists to question models; the scientific method promotes the use of a cycle of such inquiry.	• Theories are not automatically accepted and may be unpopular for long periods of time. • Philosophy and religion can be supported by arguments; science requires that theories be testable and therefore falsifiable.

2.3 Modern Atomic Theory and the Laws That Led to It • That matter is composed of atoms grew from experiments and observations. • Conceptual Connection 2.1 The Law of Conservation of Mass • Investigating the law of definite proportions requires preparing or decomposing a set of pure samples of a compound like water. • Investigating the law of multiple proportions requires preparing or decomposing sets of pure samples from related compounds like NO, NO_2, and N_2O_5.	• Measurements to establish early atomic theories were performed at the macroscopic level. The scientists observed properties for which they could collect data (e.g., mass or volume).

Chapter 2. Atoms and Elements

Lecture Outline

<table>
<tr><th><u>Terms, Concepts, Relationships, Skills</u></th><th><u>Figures, Tables, and Solved Examples</u></th></tr>
<tr><td>

2.4 Atomic Structure

- Thomson’s cathode-ray tube experiments
  - High voltage produced a stream of particles that travelled in straight lines.
  - Each particle possessed a negative charge.
  - Thomson measured the charge-to-mass ratio of the electron.
- Millikan’s oil-drop experiments
  - Oil droplets received charge from ionizing radiation.
  - Charged droplets were suspended in an electric field.
  - The mass and charge of each oil drop was used to calculate the mass and charge of a single electron.
- Thomson’s plum-pudding model: negatively charged electrons in a sea of positive charge
- Radioactivity: alpha decay provides the alpha particles for Rutherford’s experiment.
- Rutherford’s experiment
  - Alpha particles directed at a thin gold film deflect in all directions, including back at the alpha source.
  - Only a concentrated positive charge could cause the alpha particles to bounce back.
- Rutherford’s nuclear theory
  - Most mass and all positive charge contained in a small nucleus.
  - Most of atom by volume is empty space.
  - Protons: positively charged particles.
  - Neutral particles with substantial mass also in nucleus.
- Atomic number (number of protons): defining characteristic of an element
- Isotope: same element, different mass
- Properties of subatomic particles
  - atomic mass units (u)
    - Proton, neutron: ~1 u
    - Electron: ~0.006 u
  - charge
    - Relative value: –1 for electron, +1 for proton
    - Absolute value: 1.6×10^{-19} C
- Ion: atom with nonzero charge
  - Anion: negatively charged (more electrons)
  - Cation: positively charged (fewer electrons)

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- Figure 2.3: Cathode Ray Tube
- Unnumbered figure: properties of electrical charge
- Figure 2.4: Thomson’s Measurement of the Charge-to-Mass Ratio of the Electron
- Figure 2.5: Millikan’s Measurement of the Electron's Charge
- Unnumbered figure: plum-pudding model
- Figure 2.6: Rutherford’s Gold Foil Experiment
- Unnumbered figure: Rutherford Museum, McGill University
- Unnumbered figure: photo of Ernest Rutherford
- Figure 2.7: The Nuclear Atom
- Unnumbered figure: baseball
- Table 2.1: Characteristics of Subatomic Particles
- Example 2.2: Atomic Numbers, Mass Numbers, and Isotope Symbols
- Chemistry in Your Day: Where Did Elements Come From?
- Unnumbered figure: Image of gaseous pillar in a star-forming region of the Eagle Nebula

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</table>

Chapter 2. Atoms and Elements

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.4 Atomic Structure • Review the attraction, repulsion, and additivity of charges. • Discuss the physics of electric fields generated by metal plates. • A demonstration of a cathode ray tube will help students better understand Thomson's experiments. • Demonstrate how Millikan's calculation works and why he could determine the charge of a single electron. • It may be useful to give a brief description of radioactivity. Rutherford's experiment makes more sense if one knows some properties of the alpha particle and from where it comes. • Thomson identified electrons and surmised the existence of positive charge necessary to form a neutral atom. The plum-pudding model is the simplest way to account for the observations. • The analogy of the baseball and a grain of rice to a proton and an electron is meant to illustrate the difference in mass but not size. • Electrical charge can be demonstrated with static electricity. Two balloons charged with wool or human hair will repel each other. • Names of elements come from various sources. Tom Lehrer's "Element Song" can be found on the Internet. • Isotopic abundances are invariant in typical lab-sized samples because of such large numbers of atoms. • Conceptual Connection 2.2: The Nuclear Atom, Isotopes, and Ions • The history of chemistry involves considerable cultural and gender diversity. Examples include both Lavoisiers (French), Dalton (English), Thomson (English), Marie Curie (Polish/French), Mendeleev (Russian), Millikan (American), Robert Boyle (Irish), Amedeo Avogadro (Italian). • The Chemistry in Your Day box gives a broad description of the origin of atoms.	• Millikan did not measure the charge of a single electron; he measured the charge of a number of electrons and deduced the charge of a single electron. • Students often don't understand the *source* of alpha particles in Rutherford's experiments. • Students sometimes confuse the mass number as being equal to the number of neutrons, not the number of neutrons plus the number of protons. • Students logically (but mistakenly) presume that the mass of an isotope is equal to the sum of the masses of the protons and neutrons in that isotope.

Chapter 2. Atoms and Elements

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples
2.5 Atomic Mass: The Average Mass of an Element's Atom • Average atomic mass is based on natural abundance and isotopic masses. • Mass spectrometry o Atoms converted to ions and deflected by magnetic fields to separate by mass o Output data: relative mass vs. relative abundance	• Unnumbered figure: periodic table box for chlorine • Example 2.3: Atomic Mass • Figure 2.8: The Mass Spectrometer • Figure 2.9: Mass Spectrum of Xenon

2.6 Molar Mass: Counting Atoms by Weighing Them • Mole concept and Avogadro's number • Converting between moles and number of atoms • Converting between mass and number of moles	• Unnumbered figure: relative sizes of Al, C, He • Unnumbered figure: balance with marbles and peas • Example 2.4: The Mole Concept: Converting from Mass to Moles and Number of Atoms

Chapter 2. Atoms and Elements

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.5 Atomic Mass: The Average Mass of an Element's Atom • The masses of isotopes must be reconciled with an element having only whole number quantities of protons and neutrons; the values should be nearly integral since the mass of electrons is so small. • Mass spectrometry is an effective way to demonstrate where values of natural abundance are obtained.	• Students are tempted to calculate average atomic mass by adding together isotopic masses and dividing by the number of isotopes. • Atomic mass on the periodic table is usually not integral even though elements have only whole numbers of protons and neutrons.

2.6 Molar Mass: Counting Atoms and Weighing Them • Review the strategy for solving numerical problems: sort, strategize, solve, check. • Estimating answers is an important skill; the number of atoms will be very large (i.e., some large power of ten) even from a small mass or small number of moles. • Conceptual Connection 2.3: The Mole	• Many students are intimidated by estimating answers in calculations involving powers of ten.

Chapter 2. Atoms and Elements

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples

2.7 The Periodic Table of the Elements
• Periodic law and the periodic table
  o Generally arranged by ascending mass
  o Recurring, periodic properties; elements with similar properties arranged into columns: groups (or families)
• Major divisions of the periodic table
  o Metals, nonmetals, metalloids
  o Main-group elements, transition elements
• Groups (families)
  o Noble gases (group 18)
  o Alkali metals (group 1)
  o Alkaline earth metals (group 2)
  o Halogens (group 17)
• Ions with predictable charges: based on stability of noble gas electron count
  o Group 1: 1+
  o Group 2: 2+
  o Group 13: 3+
  o Group 15: 3–
  o Group 16: 2–
  o Group 17: 1–
	• Unnumbered figure: stamp featuring Dmitri Mendeleev
	• Unnumbered figure: Mendeleev's periodic table from 1871
	• Figure 2.10: Metals, Nonmetals, and Metalloids
	• Figure 2.11: The Periodic Table: Main-Group and Transition Elements
	• Unnumbered figure: the alkali metals
	• Unnumbered figure: the halogens
	• Figure 2.12: Elements That Form Ions with Predictable Charges

Chapter 2. Atoms and Elements

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.7 The Periodic Table of the Elements • Other displays of the periodic table can be found in journals (Schwartz, *J. Chem. Educ*. **2006**, *83*, 849; Moore, *J. Chem. Educ*. **2003**, *80*, 847; Bouma, *J. Chem. Educ.* **1989**, *66*, 741), books, and on the Internet. • Periodic tables are arranged according to the periodic law but can compare many features, e.g., phases of matter, sizes of atoms, and common ions. These are presented as a series of figures in the text.	• The periodic table is better at predicting microscopic properties, though macroscopic properties are also often illustrated.

Additional Problem for Converting between Number of Moles and Number of Atoms (Example 2.4)	Calculate the number of moles of iron in a sample that has 3.83×10^{23} atoms of iron.
Sort You are given a number of iron atoms and asked to find the amount of iron in moles.	**Given** 3.83×10^{23} Fe atoms **Find** mol Fe
Strategize Convert between number of atoms and number of moles using Avogadro's number.	**Conceptual Plan** atoms → mol $\dfrac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}}$ **Relationships Used** $6.022 \times 10^{23} = 1$ mol (Avogadro's number)
Solve Follow the conceptual plan. Begin with 3.83×10^{23} Fe atoms and multiply by the ratio that equates moles and Avogadro's number.	**Solution** $3.83 \times 10^{23} \text{ Fe atoms} \times \dfrac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}} = 0.636 \text{ mol Fe}$
Check	The sample was smaller than Avogadro's number, so the answer should be a fraction of a mole. The value of the sample has three significant figures, and the answer is provided in that form.

Additional Problem for Converting between Mass and Number of Moles (Example 2.4)	Calculate the number of grams of silver in an American Silver Eagle coin that contains 0.288 moles of silver.
Sort You are given the amount of silver in moles and asked to find the mass of silver.	**Given** 0.288 mol Ag **Find** g Ag
Strategize Convert amount (in moles) to mass using the molar mass of the element.	**Conceptual Plan** mol Ag → g Ag $\frac{107.87 \text{ g Ag}}{1 \text{ mol Ag}}$ **Relationships Used** 107.87 g Ag = 1 mol Ag
Solve Follow the conceptual plan to solve the problem. Start with 0.288 mol, the given number, and multiply by the molar mass of silver.	**Solution** $0.288 \cancel{\text{ mol Ag}} \times \frac{107.87 \text{ g Ag}}{1 \cancel{\text{ mol Ag}}} = 31.07 \text{ g Ag}$ 31.07 g = 31.1 g Ag
Check	The magnitude of the answer makes sense since we started with an amount smaller than a mole. The molar amount and answer both have three significant figures.

Additional Problem for the Mole Concept—Converting Between Mass and Number of Atoms (Example 2.4)	What mass of iron (in grams) contains 1.20×10^{22} atoms of Fe? A paperclip contains about that number of iron atoms.
Sort You are given a number of iron atoms and asked to find the mass of Fe.	**Given** 1.20×10^{22} Fe atoms **Find** g Fe
Strategize Convert the number of Fe atoms to moles using Avogadro's number. Then convert moles Fe into grams of iron using the molar mass of Fe.	**Conceptual Plan** Fe atoms → mol Fe → g Fe $\frac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}}$ $\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}$ **Relationships Used** $6.022 \times 10^{23} = 1$ mol (Avogadro's number) 55.85 g Fe = 1 mol Fe
Solve Follow the conceptual plan to solve the problem. Begin with 1.20×10^{22} atoms of Fe, multiply by the ratio derived from Avogadro's number, and finally multiply by the atomic mass of Fe.	**Solution** $1.20 \times 10^{22} \cancel{\text{Fe atoms}} \times \frac{1 \cancel{\text{mol Fe}}}{6.022 \times 10^{23} \cancel{\text{Fe atoms}}} \times \frac{55.85 \text{ g Fe}}{1 \cancel{\text{mol Fe}}}$ $= 1.11$ g Fe
Check	The units and magnitude of the answer make sense. The sample is smaller than a mole. The number of atoms and mass both have three significant figures.

Additional Problem for the Mole Concept (Example 2.4)	An iron sphere contains 8.55×10^{22} iron atoms. What is the radius of the sphere in centimetres? The density of iron is $7.87\ \text{g cm}^{-3}$.
Sort You are given the number of iron atoms in a sphere and the density of iron. You are asked to find the radius of the sphere.	**Given** 8.55×10^{22} Fe atoms $d = 7.87\ \text{g cm}^{-3}$ **Find** radius (r) of a sphere
Strategize The critical parts of this problem are density, which relates mass to volume, and the mole, which relates number of atoms to mass: (1) Convert from the number of atoms to the number of moles using Avogadro's number (2) Convert from the number of moles to the number of grams using the molar mass of iron (3) Convert from mass to volume using the density of iron (4) Find the radius using the formula for the volume of a sphere	**Conceptual Plan** Fe atoms → mol Fe → g Fe → V (cm^3) $\frac{1\ \text{mol Fe}}{6.022\times10^{23}\ \text{Fe atoms}}$ $\frac{55.85\ \text{g Fe}}{1\ \text{mol Fe}}$ $\frac{1\ \text{cm}^3}{7.87\ \text{g Fe}}$ V (cm^3) → r (cm) $V = \frac{4}{3}\pi r^3$ **Relationships Used** 6.022×10^{23} = 1 mol (Avogadro's number) 55.85 g Fe = 1 mol Fe d (density of Fe) = $7.87\ \text{g cm}^{-3}$ $V = 4/3\ \pi r^3$ [volume of a sphere with a radius of r]
Solve Follow the conceptual plan to solve the problem. Begin with 8.55×10^{22} Fe atoms and convert to moles, then to grams, and finally to a volume in cm^3. Solve for the radius using the rearranged equation.	**Solution** $8.55\times10^{22}\ \cancel{\text{atoms}} \times \frac{1\ \cancel{\text{mol Fe}}}{6.022\times10^{23}\ \cancel{\text{atoms}}} \times \frac{55.85\ \cancel{\text{g Fe}}}{1\ \cancel{\text{mol Fe}}} \times \frac{1\ \text{cm}^3}{7.87\ \cancel{\text{g Fe}}} = 1.00757\ \text{cm}^3$ $r = \sqrt[3]{\frac{3\ V}{4\ \pi}} = \sqrt[3]{\frac{3\times1.00757\ \text{cm}^3}{4\ \pi}} = 0.622\ \text{cm}$
Check	The units (cm) are correct and the magnitude of the answer makes sense compared with previous problems.