CH 221 Chapter Two Concept Guide

1. Origins of Atomic Theory

The simple picture of atoms as tiny spheres in constant motion goes back a long way in history - to the Greek philosopher Levucippus and his student, Democritus (460 - 370 BC). Democritus reasoned that if a bit of matter was divided into smaller and smaller pieces, one would ultimately arrive at a tiny particle that could not be further divided. He described this particle with the word "atom," which means uncuttable.

Democritus used his concept of atoms to explain physical properties and actions, such as density, hardness, and evaporation. The theory was untestable and unsupported, however, and remained so for more than 2000 years.

By the time a useful theory of atoms was developed early in the 19th century, much progress had been made in science and chemistry. The existence of gases was recognized and widely studied. Some basic physical laws, such as the conservation of matter, had been elucidated. Mathematics was recognizes as a key to understanding the universe. The value of experimentation and measurement was established. Science had evolved into a form we can recognize today.

The most significant progress was made in a concentrated period near the end of the 19th and at the beginning of the 20th centuries. By the dawn of the 19th century, certain basic rules or scientific laws had been formulated and were widely accepted: the law of conservation of matter, the law of constant composition, and the law of multiple proportions. To be accepted, any theory of matter must satisfy these laws.

In 1803, John Dalton revived the idea of atoms as real objects having size and mass, and he developed the first useful atomic theory. He linked the existence of elements, which cannot be chemically decomposed, to the idea of atoms as invisible units.

Dalton's postulates were generally accepted at the time because they were useful and violated none of the established scientific laws.

Another step in understanding atomic structure was the discovery of radioactivity, which helped determine that atoms can in fact break down, and by implication, have a structure we can analyze. Marie Curie (1867 - 1934) shared the Nobel Prize in Physics in 1903 with her husband, Pierre Curie (1859 - 1906) and Henry Becquerel (1852 - 1908) for discovering the phenomenon of "radioactivity," a process in which atoms of naturally radiating substances emit their unusual rays as they spontaneously disintegrate.

2. Electrons

The idea of electrons existed long before the particles were found to be a part of atoms. The English scientist Michael Faraday (1791-1867) performed many experiments with electricity, which led to the concept of a fundamental particle of electricity (then believed to be distinct from atoms). The name "electron" was suggested in 1891, and in 1897, J. J. Thomson of Cambridge University in England proved the particle nature of electrons in an experiment using cathode rays.

Robert Andrews Millikan (1868-1953) followed Thomson, devising an elegant, now classic experiment to measure the charge of the electron using oil droplets with negative charges, now called the "Millikan Oil Drop Experiment." Millikan's measurements of charges on oil droplets led to the discovery of the charge on the electron. Since the charge-to-mass ratio of the electron was known, its mass could be calculated. The currently accepted value for the mass is 9.109389×10^{-28} g, and that for the charge is $1.60217733 \times 10^{-19}$ C, where C, the Coulomb, is the SI unit of electric charge.

3. Protons

It was a student of Thomson's, Ernest Rutherford, who, in 1914, discovered the proton. Rutherford focused on the particles of canal rays, those that flow in the opposite direction to cathode rays in gas-discharge tubes. He found that the lightest and simplest canal-ray particles are formed when a gas-discharge tube contains hydrogen. Eventually he showed that particles identical to hydrogen atoms with one electron missing, H^+, are present in all matter. He named these particles protons, a fundamental subatomic particle with a positive charge equal in magnitude to the negative charge of the electron.

4. Neutrons

Because they have no charge, neutrons are much harder to detect than protons and electrons, and the search for the neutron lasted many years. It began in 1920 with Rutherford's idea that the nucleus might contain an uncharged particle with a mass close to that of the hydrogen atom. James Chadwick, a member of Rutherford's laboratory in Cambridge, saw the answer in alpha-particle experiments reported by other workers, in which a "highly penetrating" radiation from beryllium knocked protons out of paraffin with great force. Chadwick believed this radiation was a beam of uncharged particles, each with the mass expected for a neutron. After performing his own experiments, Chadwick proved conclusively the existence of the neutron in 1932.

5. The Modern View of an Atom

The existence of protons and electrons led to what became known as the Thomson model of the atom. Thomson proposed that an atom consists of a positively charged, uniform sphere of relatively large volume and low density, in which negatively charged electrons are embedded. A single experiment, however, destroyed this model.

Rutherford's now famous experiment involved bombarding a thin gold foil with positively charged alpha particles.

If Thomson's model were correct, the positively charged alpha particles would have plunged through the atoms of the foil like bullets through jelly. Instead, some of the particles were deflected backward. Obviously, they were hitting something very dense and very small. Rutherford and his group had discovered the nucleus of the atom.

6. Basics of Matter

All objects have physical properties, such as color, boiling point, melting point, magnetism, and viscosity. Physical properties are observed and measured without changing the composition of a substance.

a. Atoms, Molecules, Elements, and Compounds

All matter is constructed of building blocks called atoms. These particles are the simplest form of matter and the smallest that retain the chemical properties of that element.

Some pure substances, such as carbon, oxygen, and hydrogen, are composed of only one type of atom and are classified as elements. Atoms can link together to form bigger building blocks, or molecules. When two or more different kind of atoms combine, the combination is called a chemical compound. More specifically, the definition of a compound is limited to "pure" substances, the atoms of which are in a fixed ratio. For example, water is a compound that has two hydrogen atoms for each oxygen atom.

A molecule is the smallest entity that retains the chemical properties of the compound. Each molecule has a definite number of atoms, represented by the compound's formula. Letters in a chemical formula are symbols for the element; numbers used as subscripts indicate the number of atoms of that element in a single molecule.

Problem
Name the element represented by each of the following symbols:
(a) As (b) Be (c) B (d) V (e) Tl

Solution
(a) Arsenic (b) Beryllium (c) Boron (d) Vanadium (e) Thallium

Problem
Write the symbol for each of the following elements:
(a) xenon (b) magnesium (c) cobalt (d) copper (e) lead (f) silver (g) gold.

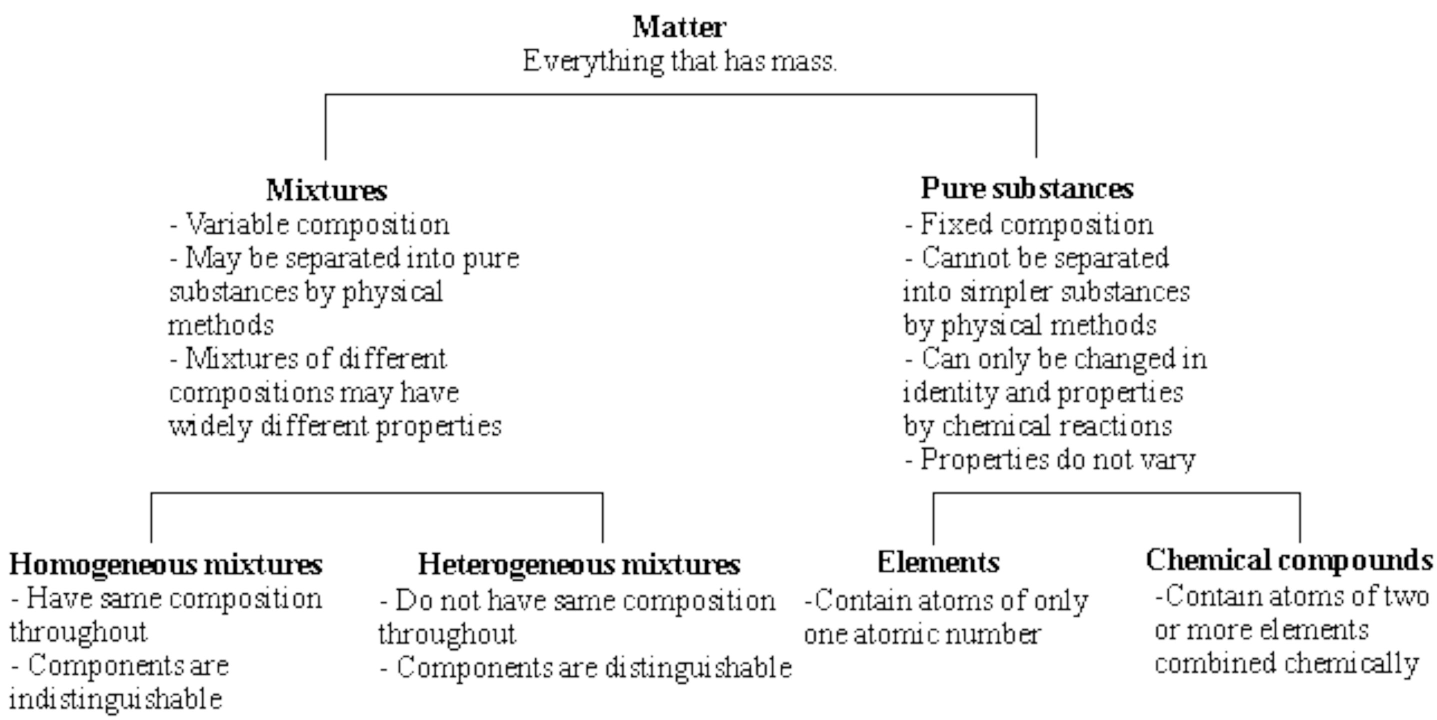
Solution
(a) Xe (b) Mg (c) Co (d) Cu (e) Pb (f) Ag (g) Au

b. States of Matter

An easily observed property of matter is its physical state, or phase. Almost all substances are solids, liquids, or gases, and virtually all matter is found in the solid state at low temperatures. At higher temperatures, solids generally melt to form liquids. Further heating may cause liquids to evaporate to form gases.

c. Mixtures and Pure Substances

All matter can be classified as a pure substance or a mixture by examining its properties and composition. Typically, there is further classification of mixtures into a homogeneous mixture or heterogeneous mixture, and of pure substances into elements and chemical compounds. For example, air is a homogeneous mixture, whereas dirt is a heterogeneous mixture because the components remain physically separate and can be seen as separate components. The general classification of matter is summarized in the diagram below.



Reference: Bailar, JC; Moeller T; Kleinberg J; Guss CO; Castellion ME; Metz C. Chemistry, 3rd ed. San Diego: Harcourt Brace Jovanovich, 1989.

Problem: Using the "Classification of Matter" terminology given in this lesson, describe each of the following and identify the number of phases in each sample:

(a) a drop of mercury (b) an ice cube (c) a melting ice cube (d) a puddle of water

Solution

(a) matter, pure substance, element; 1 phase
(b) matter, pure substance, compound; 1 phase
(c) matter, heterogeneous mixture; 2 phases
(d) matter, pure substance, compound; 1 phase

7. Atomic Mass

Question

Atoms frequently gain or lose electrons during chemical processes. Does this substantially affect their mass?

Approach

Select an element, add or subtract electrons, and calculate its new mass. Compare this new mass to the mass of the element.

Solution

As an example, let's look at an iron atom with a nucleus containing 26 protons and 30 neutrons having a total mass of 9.28836×10^{-23} g. If iron loses 2 electrons (mass of electron = 9.109389×10^{-28} g) to form Fe^{2+}, its mass is 9.28818×10^{-23} g.

$$9.28836 \times 10^{-23} - (2)(9.109389 \times 10^{-28} \text{ g}) = 9.28818 \times 10^{-23} \text{ g}$$

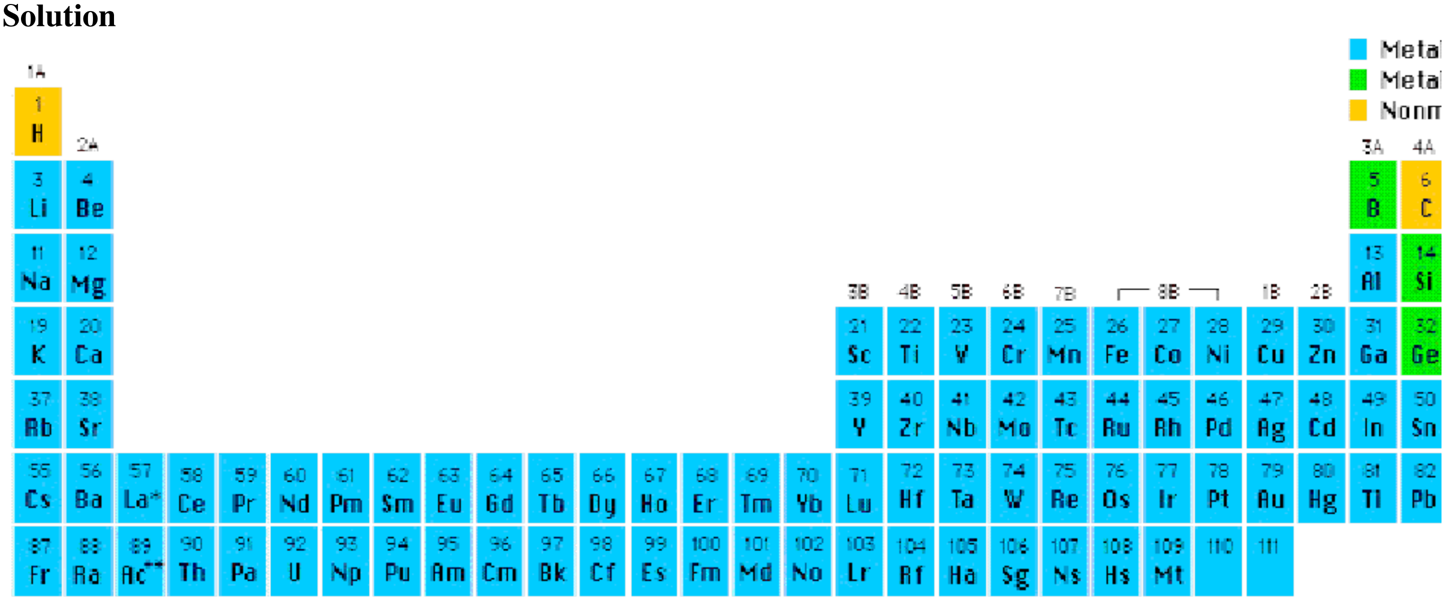
Based on this calculation, it is clear that the loss of electrons does not significantly change the mass of the element. The mass of Fe and Fe^{2+} is approximately the same, with a difference of only 1.8×10^{-27} g, or 0.002%.

8. Examining the Periodic Table

Question

What would the periodic table look like if the Lanthanide and Actinide series were not in a separate section?

Solution



9. Element Characteristics

Problem
List five attributes of chlorine.

Approach
Consider where chlorine is in the periodic table.

Solution
Chlorine is:

1. in Group 7A	2. a halogen
3. a non-metal	4. in Period 3
5. a Main Group element	6. a gas
7. very reactive	8. an element that reacts violently with alkali metals to form salts

10. Element Characteristics

Problem
List five attributes of sodium.

Approach
Consider where sodium is on the periodic table.

Solution
Sodium is:

1. an alkali metal	2. in Group 1A
3. a Main Group element	4. in Period 3
5. a solid	6. a conductor of electricity
7. malleable and ductile	8. reactive with water to produce hydrogen gas and an alkaline solution (NaOH)
9. not found free in nature	10. always found combined with other elements as compounds

11. Mendeleev's Periodic Table

Question

Examine Mendeleev's periodic table (circa 1872).

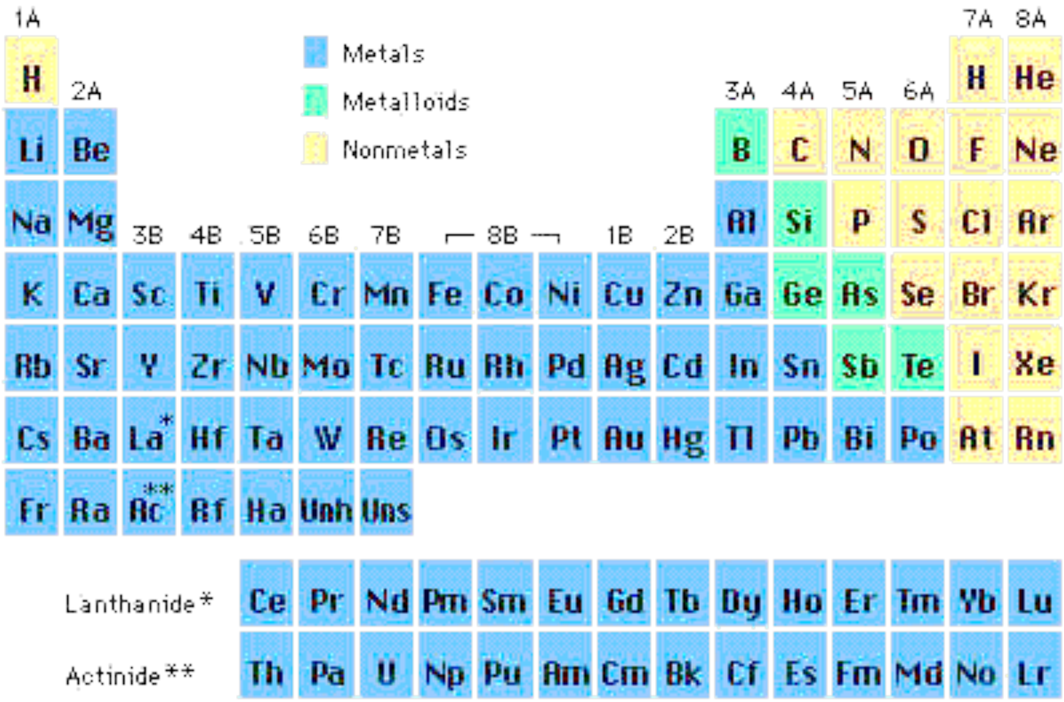
TABELLE II

REIHEN	GRUPPE I. — R^2O	GRUPPE II. — RO	GRUPPE III. — R^2O^3	GRUPPE IV. RH^4 RO^2	GRUPPE V. RH^3 R^2O^5	GRUPPE VI. RH^2 RO^3	GRUPPE VII. RH R^2O^7	GRUPPE VIII. — RO^4
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	—=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	—=100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

To what elements on the modern periodic table do the missing elements having masses of 68 and 70 grams per mole correspond? How do these predicted masses on Mendeleev's periodic table compare to those on the modern periodic table?

Solution

On Mendeleev's periodic table, the element having a mass of 68 refers to Gallium and that having a mass of 70 refers to Germanium. This is plausible because these elements fall between zinc and arsenic, as they do on Mendeleev's periodic table. In addition, the mass of Gallium is 69.72 grams per mole, which is only a difference of +1.72 grams per mole compared to Mendeleev's estimated 68 grams per mole. With regard to Germanium, the mass of this element is 72.59 grams per mole, a difference of only 2.59 grams per mole from Mendeleev's approximation. This is remarkable considering that Mendeleev did not observe these elements, but predicted their existence and properties by inference of other known elements and the structure of his table.



12. Molar Mass

Question

What is the molar mass of NaOH?

Approach

To find the molar mass of a compound, we must list the elements in the compound's formula and determine the number of atoms of each element in the formula. Then, for each element, we need to look up the molar mass. Once we have found the mass contributed by each element for one mole of the compound, the molar mass is calculated by adding these individual masses.

Solution

1 mol Na in NaOH = (1)22.98977 g/mol = 22.98977 g
1 mol O in NaOH = (1)15.9994 g/mol = 15.9994 g
1 mol H in NaOH = (1)1.0079 g/mol = 1.0079 g

Molar mass of 1 mol NaOH = 22.98977 g + 15.9994 g + 1.0079 g = 39.9971 g/mol NaOH

13. Molar Mass

Question

What is the molar mass of $Cu(NO_3)_2$?

Approach

To find the molar mass of a compound, we must list the elements in the compound's formula and determine the number of atoms of each element in the formula. Then, for each element, we need to look up the molar mass. Once we have found the mass contributed by each element for one mole of the compound, the molar mass is calculated by adding these individual masses.

Solution

1 mol Cu in $Cu(NO_3)_2$ = (1)63.546 g/mol = 63.546 g
2 mol N in $Cu(NO_3)_2$ = (2)14.0067 g/mol = 28.0134 g
6 mol O in $Cu(NO_3)_2$ = (6)15.9994 g/mol = 95.9964 g

Molar mass of 1 mol $Cu(NO_3)_2$ = 63.546 g + 28.0134 g + 95.9964 g = 187.556 g/mol of $Cu(NO_3)_2$

14. Converting Mass to Moles

Question
What quantity, in moles, does 107 g HBr represent?

Solution
The molar mass of HBr is 80.912 g/mol. The number of moles of HBr is:
107 g HBr * 1 mol HBr/80.912 g HBr = 1.32 mol HBr

15. Molar Mass and Moles

Question
Which represents a greater number of moles: 4.5 g of carbon dioxide or 4.5 g of sodium chloride?

Solution
The molar masses are 44.01 g for CO_2 and 58.44 g for NaCl. The numbers of moles of each compound is calculated using the molar mass:
4.5 g CO_2 * 1mol CO_2/44.01 g CO_2 = 0.10 mol CO_2
4.5 g NaCl * 1 mol NaCl/58.44 g NaCl = 0.077 mol NaCl

There are more moles of CO_2 than of NaCl.

16. Atomic Mass

Question
What is the average atomic mass of chlorine?

Mass of ^{35}Cl	34.96885 amu
Mass of ^{37}Cl	36.96712 amu
Isotopic abundance of ^{35}Cl	75.77 %
Isotopic abundance of ^{37}Cl	24.23 %

Approach
Consider the mass and abundance of the isotopes of chlorine.

Solution:
Step 1. There are 2 naturally occurring isotopes of chlorine: ^{35}Cl and ^{37}Cl. A sample of chlorine shows that the 2 isotopes are not present in equal amounts. The percent abundance is calculated by dividing the number of atoms of a given isotope by the total number of atoms of all isotopes of that element, times one hundred.

$$\text{Percent abundance} = \frac{\text{number of atoms of a given isotope}}{\text{total number of atoms of that element}} \times 100\%$$

Step 2. The average atomic mass is

$$\text{Atomic mass} = \sum_{\text{all isotopes}} \text{isotopic mass x fractional abundance}$$

For chlorine, the atomic mass is

Atomic mass = 34.96885 amu * 0.7577 + 36.96712 amu * 0.2423 = 35.45 amu

17. Converting Mass to Atoms

Question

The 1989 nutritional recommended dietary allowance (RDA) of iron for a female age 19-24 is 15 mg. How many iron atoms is this?

Approach

Mass must be converted to moles, then moles must be converted to atoms. Avogadro's number (6.022×10^{23}) will be needed, as will the molar mass of iron. Note: in molar calculations, all masses must be converted to grams.

Solution:

Follow the series of multiplication steps below to convert mass to atoms.

$$(15 \text{ mg Fe}) \left(\frac{1 \text{ g}}{1000 \text{ mg}}\right)\left(\frac{1 \text{ mol Fe}}{55.847 \text{ g Fe}}\right)\left(\frac{6.022 \text{x} 10^{23}}{1 \text{ mol Fe}}\right) = 1.6 \times 10^{20} \text{ Fe atoms}$$

In 15 mg of iron there are 1.6×10^{20} Fe atoms.

18. Converting Mass to Molecules and Atoms

Question

How many ozone molecules and how many oxygen atoms are contained in 48.00 g ozone, O_3?

Approach

First, write the equation for the synthesis of ozone. Then, to calculate the number of ozone molecules in 48.00 g ozone, convert mass to molecules using the molecular weight of ozone and Avogadro's number (6.022×10^{23}). Next, to calculate the number of oxygen atoms, start with the mass of ozone, convert this to grams, use a mole-to-mole ratio to convert ozone to O_2, and finally, use Avogadro's number to obtain the number of oxygen atoms.

Solution:

Step 1. Write the equation for the synthesis of ozone, and calculate its mass.

$3 O_2(g) \rightarrow 2 O_3(g)$ $\qquad$ $O_3 = 3(15.999 \text{ g/mol}) = 47.997 \text{ g/mol}$

Step 2. To calculate the number of ozone molecules from its mass, multiply grams of ozone, molecular weight of ozone, and Avogadro's number.

$$48.00\text{g } O_3 \cdot \frac{1 \text{ mol } O_3}{49.997\text{g } O_3} \cdot \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol } O_3} = 6.022 \times 10^{23} \text{ molecules } O_3$$

The number of molecules of ozone in 48.00 grams is 6.022×10^{23} molecules.

Step 3. To calculate the number of atoms of oxygen, multiply three atoms of oxygen per molecule of ozone and the number of molecules of ozone.

$$6.022 \times 10^{23} \text{ molecules } O_3 \cdot \frac{3 \text{ atoms O}}{1 \text{ molecule } O_3} = 1.807 \times 10^{24} \text{ atoms O}$$

The number of oxygen atoms in 48.00 g of ozone is 1.807×10^{24} atoms.