MCGRAW-HILL RYERSON CHEMISTRY 11 TEACHER'S RESOURCE

ONTARIO EDITION

UNIT 2

Quantities in Chemical Reactions

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Notes

CHAPTER 5 CURRICULUM EXPECTATIONS

Expectations	Text Section/ThoughtLab/ExpressLab/Investigation
Overall Expectations ■ QCR V.01 demonstrate an understanding of the mole concept and its significance in the analysis of chemical systems	■ 5.2 The Avogadro Constant and the Mole, pp. 161–169 ■ ThoughtLab: The Magnitude of the Avogadro Constant, p. 165 ■ 5.3 Molar Mass, pp. 170–182 ■ Investigation 5-A: Modelling Mole and Mass Relationships, pp. 172–173
Specific Expectations *Understanding Basic Concepts* ■ QCR 1.01 demonstrate an understanding of Avogadro's number, the mole concept, and the relationship between the mole and molar mass	■ 5.2 The Avogadro Constant and the Mole, pp. 161–169 ■ ThoughtLab: The Magnitude of the Avogadro Constant, p.165 ■ 5.3 Molar Mass, pp. 170–182 ■ Investigation 5-A: Modelling Mole and Mass Relationships, pp. 172–173
■ QCR 1.02 explain the relationship between isotopic abundance and relative atomic mass	■ 5.1 Isotopes and Average Atomic Mass, pp. 152–160 ■ ExpressLab: A Penny for Your Isotopes, p. 10
■ QCR 1.05 state the quantitative relationships expressed in a chemical equation (e.g., in moles, grams, atoms, ions, or molecules)	■ 5.3 Molar Mass, pp. 170–182 ■ Investigation 5-A: Modelling Mole and Mass Relationships, pp. 172–173
Developing Skills of Inquiry and Communication ■ QCR 2.01 use appropriate scientific vocabulary to communicate ideas related to chemical calculations (e.g., stoichiometry, percentage yield, limiting reagent, mole, atomic mass)	■ 5.1 Isotopes and Average Atomic Mass, pp. 152–160 ■ 5.2 The Avogadro Constant and the Mole, pp. 161–169 ■ 5.3 Molar Mass, pp. 170–182
■ QCR 2.03 solve problems involving quantity in moles, number of particles, and mass	■ 5.2 The Avogadro Constant and the Mole, pp. 161–169 ■ ThoughtLab: The Magnitude of the Avogadro Constant, p. 165 ■ 5.3 Molar Mass, pp. 170–182 ■ Investigation 5-A: Modelling Mole and Mass Relationships, pp. 172–173
■ QCR 2.07 calculate, for any given reactant or product in a chemical equation, the corresponding mass or quantity in moles or molecules of any other reactant or product	■ 5.3 Molar Mass, pp. 170–182
Relating Science to Technology, Society, and the Environment ■ QCR 3.01 give examples of the application of chemical quantities and calculations (e.g., in cooking recipes, in industrial reactions, in prescription drug dosages)	■ 5.1 Isotopes and Average Atomic Mass, pp. 152–160 ■ 5.2 The Avogadro Constant and the Mole, pp. 161–169 ■ 5.3 Molar Mass, pp. 170–182

CHAPTER 5

Counting Atoms and Molecules: The Mole

Student Textbook pages 152–160

Contents

Key Terms

average atomic mass, p. 154
Avogadro constant, p. 162
isotopic abundance, p. 153
mass spectrometer, p. 154
mole, p. 162
molar mass, p. 170
weighted average, p. 155

Introduction

The notion of the atom dates back almost 2500 years to 450 B.C. Greek philosophers Leucippus and Democritus believed that all matter could be subdivided into indivisible units called atomia (or atoms). In the early 1800s, John Dalton developed his own theory of atoms and what we now know as the beginning of the modern atomic theory. Dalton stated that atoms are the smallest form for which matter can exist. Dalton's theory also went on to describe how atoms from different elements have different properties and that these elements can combine in small, whole number ratios to form compounds. The atom and the theories surrounding it are the foundations for chemistry, but the question still remains, since atoms are so small, how can scientists work with matter and understand chemical phenomena at an atomic scale? The information contained within Chapter 5 of the student textbook begins to answer these questions. The masses of individual atoms can be determined through the process of mass spectrometry. These masses are then used to determine the average atomic mass of each element. Since most elements have isotopes, it is necessary to know the mass and relative abundance of each individual isotope to obtain an average mass for the atoms of the element. Avogadro's constant (N_A = 6.02 x 10^{23}) is used to develop a relationship between the average atomic mass (expressed in amu) and useful quantities of measurements (the mole, molar mass) that can be used by chemists in a laboratory setting.

5.1 Isotopes and Average Atomic Masses

Student Textbook pages 152–160

This section builds on students' prior knowledge of relative atomic mass and isotopes. The idea of weighted averages is introduced. Then the idea of isotopic abundance is linked to average atomic mass. The principle of the mass spectrometer is outlined and its uses in determining isotopic abundance are explained.

Relating Atomic Masses to Macroscopic Masses

Student Textbook pages 152–153

Science Background

Prior to this unit, students have been presented with the idea that atoms are composed of protons, neutrons and electrons. Of these three subatomic particles, it is only the protons and neutrons that contribute significantly to the mass of an atom.

Mass Number = number of protons + number of neutrons

Subatomic Particle	Mass (kg)	Mass (u)
proton	1.67 x 10^{-27}	1.0073
neutron	1.67 x 10^{-27}	1.0087
electron	9.02 x 10^{-31}	0.0005486

GETTING STARTED!

Place a large glass jar of jellybeans (try to put only three or four different colours of jellybeans in the jar) on your desk and challenge the students to guess how many jellybeans are in the jar. After a few random (and probably wrong) guesses at the amount, have the students devise a method to determine the number of jellybeans if they were given the following items:

- identical glass jar (empty)
- balance
- small sample of jellybeans
- empty beaker

Students will most likely come up with a method of indirect counting for determining the number of jellybeans in the jar. Typical steps for this method include:

- determining the mass of the empty jar and the full jar (subtract one from the other to determine the mass of all of the jellybeans)
- measuring the mass of a small amount of jellybeans to determine the mass of one jellybean
- dividing the mass of all of the jellybeans by the mass of one jellybean to determine the number of jellybeans in the jar

Explain to the students that this same method of counting is employed to determine the mass of entities that we cannot see. You may, however, want to confirm for your students that this method does work. In that case, once students have come up with their method, you could bring out the appropriate materials and have several students demonstrate in front of the class. Finally, divide the jelly beans up for counting to confirm the calculated value. This visual can be kept at the front of the room to explain both isotopes and weighted averages, both topics that will be explored in this chapter.

Note: Students should never eat anything that they have been working with in the laboratory.

A single carbon atom is assigned a mass of exactly 12 atomic mass units (u). Therefore one atomic mass unit has 1/12 the mass of a carbon 12 atom. Since the protons and neutrons do not have masses of exactly 1 u, the masses of atoms, other than carbon-12 are not exact integer values.

Teaching Strategies

- As a class, read over the chapter opener (page 151) and discuss the common types of measurement used in students' day-to-day activities. The purpose of the opener is to introduce the need for indirect counting of very small chemical entities, atoms, and molecules. Remind students that cooking is all about chemistry, and recipes are an important part of food preparation. Discuss the fact that North American recipes rely heavily on measured volumes of ingredients, while kitchen scales are used more in Europe, meaning ingredients are measured by mass. For example, 1 litre (or four cups) of flour in a North American recipe would be 450 grams in Europe. Whichever method is used, you get the same tasty muffins.
- Ask students how they think the concept of indirect counting applies to recipes. Explain that recipe writers have learned the proper ratio of masses or volumes of these ingredients to ensure perfect muffins. In effect, they are counting molecules indirectly using mass or volume. Then ask students how this relates to chemistry. Explain that chemists too can use different measures of the same amount of a substance. For instance, in the following reaction:

  $2H_{2(g)} + O_{2(g)}$ › $2\ H_2O_{(l)}$

  The balanced chemical equation means that two molecules of hydrogen react with one molecule of oxygen to produce two molecules of water. Dealing with individual atoms and molecules is next to impossible to do within a lab, so scientists need a way to relate large numbers of atoms or molecules to their masses
- Before dealing with isotopes quantitatively, a quick review of how to calculate the numbers of subatomic particles within an atom can be done. Students should understand the following relationships
  number of electrons = atomic number
  number of protons = atomic number
  number of neutrons = mass number – atomic number
- To give students a better understanding for the need of a relationship between the atomic scale and a macroscopic quantity, have students determine the masses of various atoms in kilograms using the above table followed by a demonstration of the minimum mass that a typical lab balance can detect.
- List the isotopes of hydrogen and their atomic masses.

Isotope Symbol	Number of Protons	Number of Neutrons	Atomic Mass (u)
$^{1}_{1}H$	1	0	1.0078
$^{2}_{1}H$	1	1	2.0141
$^{3}_{1}H$	1	2	3.0160

Have students recall that isotopes are atoms of the same element with the same atomic number but different atomic masses. Then ask why the average atomic mass, as shown in the Periodic Table, is 1.01 u. This allows the introduction of the idea of weighted averages. Explain that there are not equal numbers of these three types of hydrogen atoms in naturally occurring samples of hydrogen.

- Back to the jellybeans! Have students examine the jar of jellybeans – how many different colours of jellybeans are present in the jar. Have students imagine the jellybeans are all one type of atom. Why do the jellybeans have different colours? It could be that the different colours represent different isotopes. Just as atoms that are isotopes have different numbers of neutrons in the nucleus the isotopic jellybeans have different colours and tastes. Jellytopes.
- Bring in a number of spoons (also available from the family studies room) and have the students examine the questions posed in Figure 5.2's caption.

SUPPORTING DIVERSE STUDENT NEEDS

For students who have not yet mastered theory and calculations involving subatomic particles, Blackline Master 5-1: Subatomic Particles and Isotopes (Reinforcement) can be completed in order to give them a better understanding of isotopes and atomic structure.

[Figure to come]

Figure 5.2
Student Textbook page 153

The similarity between finding the average masses for the spoons and the isotopes is that both averages must be weighted averages, to reflect the unequal proportions of the items being averaged. Also, the spoons all belong to the category "spoon," but have three different types of spoon with three masses. This is similar to the three magnesium isotopes. They all belong to the group "magnesium," but have three different characteristic masses.

Isotopic Abundance, and Average Atomic Mass and the Periodic Table

Student Textbook pages 153–154

Science Background

Most elements exist in at least two isotopic forms in definite proportions. Therefore, weighted averages of the masses of these isotopes, called average atomic masses, are used for mass calculations for samples of elements. For example, chlorine is 75 percent Cl-35 and 25 percent Cl-37, which gives an atomic mass of:
[(0.75)(35) + (.25)(37)] u or 35.5 u.

Note that protons and neutrons do not have masses of exactly 1 u, their masses are approximately 1 u. Therefore, the mass number of an atom approximately represents its mass in atomic mass units.

To calculate these weighted averages, the relative proportions of the isotopes in naturally occurring samples must be known. These percentage abundances of isotopes can be found using mass spectrometry.

CHEM FACT
Student Textbook page 153

Magnesium deficiency has been linked to heart disease; its lack can result in clots in the heart and the brain and in calcium deposits in the kidneys and the heart.

Teaching Strategies

- For practice in determining relative amounts in isotopic abundance, pose the following problem to the class:
  In a class of 30 brilliant chemistry students, five of the students have blue eyes, 15 have brown eyes and ten have green eyes. Determine the relative amount of each eye colour found in the class and express each as a percentage.
- Bring up the problem of the spoons (Figure 5.2, page 153) and find a relationship between the mass of a large number of atoms and the mass of a single atom. How will the existence of isotopes affect the relationship between the mass of an atom and the mass of many atoms? Have students discuss the need for a weighted average.
- Use Blackline Master 5-2: Rubber Stopper Isotopes (Skill Builder) to link the concept of weighted averages to isotopes. If time allows, this activity can be performed in class as a demonstration or in small groups.

Common Misconceptions

Some students believe that the average atomic mass applies to individual atoms. Point out that this is not true any more than an average family size of 3.4 means that any specific family has a fraction of a person as a member. The same thinking also applies to atoms, as fractional values of protons, neutrons, and electrons cannot exist within a single atom.

Working with Weighted Averages, and Calculating Average Atomic Mass

Student Textbook pages 155, 159

Teaching Strategies

- One use of weighted averages involves class marks, a topic of real interest to students. Answer, with the help of the class, MindStretch (page 155) on calculating final marks.
- The MindStretch activity can be expanded so students in the class can calculate their own marks using the weighting system used for their chemistry course.
- More jellybeans – use the jar of jellybeans to help students understand the concept of weighted averages better. Say there were four colours of jellybeans in the jar: yellow, red, green, and black. Assign a "relative abundance" to each jellybean colour and a mass. Using this data, have students calculate the new average jellybean (jellyatomic) mass. Using this new mass – see how their indirect counting method of the jellybeans can possibly yield a different answer.
- In calculating weighted averages, the relative abundances of each isotope are given. In the case of silver, have students visualize a pile of 100 silver atoms. In this pile, 51.8 atoms would have a mass of 106.9 u and the other 48.2 atoms would have a mass of 108.9 u. If it is difficult to envision having .2 of a silver atom, then extend the number of silver atoms in the pile to 1000, where 518

SUPPORTING DIVERSE STUDENT NEEDS

ESL students who have difficulty understanding the concept of isotopic abundance may benefit from a visual example. Bring a number of chocolate chip cookies into class, each with different amounts of chocolate chips: plain (very few chips), regular (average amount of chips), and chocolate lovers' supreme (a lot of chips). The three cookies are all about the same size, they all taste about the same (have the same "chemical properties") but they have different number of "neutrons"—chocolate chips. These would be "isotopic" cookies—having different masses.

Gifted students can take on a bonus project on mass spectrometry based on the Tools and Techniques item on page 156 of the student textbook. If they have access to The Idea Bank Collation or the Internet (www.s17science.com), they can attempt to create a simulation of a mass spectrometer to demonstrate to the rest of the class. Alternatively, they could investigate the contributions of F.W. Aston to the mass spectrometer.

SUPPORTING DIVERSE STUDENT NEEDS

To aid ESL students in developing confidence and competence in these calculations, be sure to go over student answers to assigned practice and review problems on the board, by peer checking, or with posted answers.

For students who need extra help Blackline Master 5-3: Weighted Averages and Isotopes (Skill Builder) provides a step-by-step walkthrough in solving weighted average isotopic abundance problems. This can be used to assist those students who are having difficulties with the mathematics of this unit.

would have a mass of 106.9 u and the other 482 atoms having a mass of 108.9 u.

- When working through the practice problems on determining Average Atomic Mass (page 157) emphasize the importance of using a stepwise and logical manner for problem solving. Many students will rush through these problems and wind up with incorrect answers.
- Provide students with a starting formula, such as the one below, to aid in the process of solving Average Atomic Mass Problems.

$$\text{Average Atomic Mass} = \frac{\text{Total Mass of Isotope A} + \text{Total Mass of Isotope B} + \ldots}{\text{Total Number of Atoms}}$$

Total Mass of Isotope A = *Relative Abundance of Isotope A x Mass of Isotope A
Total Mass of Isotope B = *Relative Abundance of Isotope B x Mass of Isotope B
Total Number of Atoms = 100
*Relative abundance should be listed as a percent (out of 100)

- As shown in the Sample Problem on page 159, the relative abundance of isotopes can be determined algebraically. Once again, a problem such as this should be worked out slowly to ensure that the students are getting the point. Many times in chemistry, formulas are used to calculate a variety of quantities. By getting students to think about what they are solving for, and having them work out the solutions slowly, they should have very few problems with mathematical concepts used in the remainder of this course.
- Section Review problem 4 (page 160) poses a question about radioisotopes. For some practical uses of radioisotopes, go to the McMaster University Nuclear Reactor web site: http://www.science.mcmaster.ca/mnr/mnrhome.htm.

Tools and Techniques: The Mass Spectrometer

Student Textbook page 156

Irwin Talesnick, former professor of education at Queen's University has designed a demonstration showing the principle of the mass spectrometer. See The Idea Bank Collation: A Handbook for Science Teachers (Irwin Talesnick (Kingston, ON: S17 Science Supplies and Services Co Ltd., 1991), Idea number 51 (ISBN 0-9691743-2-2)). If your science department does not have this collation, details are available at www.s17science.com. Both the Idea Bank Collation and a kit to carry out the demonstration are available.

mind STRETCH

Student Textbook page 155

The final mark = 25(114/130) + 30(261/300) + 5(90/95) + 10(21/25) + 30(70/80) = 87.41 or 87%.

CHECKPOINT

Student Textbook page 155

Consider the isotopes of helium, $^{3}_{2}He$ and $^{4}_{2}He$.

Isotope	Atomic Number	Mass Number	Atomic Mass (u)	Percent Abundance
He-3	2	3	3.0161	0.00014
He-4	2	4	4.0026	99.99986

Each isotope has its own mass number equal to the number of nucleons (protons and neutrons) in its nucleus. The average atomic mass is a weighted average of the two atomic masses (3.0161u, 4.0026u). Since helium is almost entirely He-4, the average atomic mass is very close to 4.0026u. In this case, the average atomic mass, 4.0026u, is very close to the mass number of the dominant isotope, 4.

Solutions for Practice Problems

Student Textbook pages 157–158

Note: *aam* **= average atomic mass.**

1.

Isotope	Atomic Mass (u)	Percentage Abundance
B-10	10.01	19.78
B-11	11.01	80.22

aam = [19.78(10.01) + 80.22(11.01)] u / 100
= [197.998 +883.22] u / 100
= 10.81 u

2.

Isotope	Atomic Mass (u)	Percentage Abundance
Si-28	27.98	92.23
Si-29	28.97	4.67
Si-30	29.97	3.10

aam = [92.23(27.98) + 4.67(28.97) + 3.10(29.97)] u / 100
= [2580.59 + 135.29 + 92.91] u / 100
= 28.09 u

3.

Isotope	Atomic Mass (u)	Percentage Abundance
Cu-63	62.93	69.1
Cu-65	64.93	30.9

aam = [69.1(62.93) + 30.9(64.93)] u / 100
= [4348.5 + 2006.3] u / 100
= 63.55 u

4.

Isotope	Atomic Mass (u)	Percentage Abundance
Pb-204	204.0	1.37
Pb-206	206.0	26.26
Pb-207	207.0	20.82
Pb-208	208.0	51.55

aam = [1.37(204) + 26.26(206) + 20.82(207) + 51.55(208)] u / 100
= [279.48 + 5409.56 + 4309.74 + 1072.24] u /100
= 207.2 u

CHEM FACT

Student Textbook page 157

Subatomic Particle	Mass (u)
proton	1.0073
neutron	1.0087
electron	0.0005486

Since the subatomic particle masses are not whole numbers, they seldom add up to a whole number. As an example, consider $^{1}_{3}Li$ with 3 p, 4 n, and 3 e.

Atomic mass = 3(1.0073)
+ 4(1.0087)
+ 3(0.0005486)
= 7.0583 u

CHECKPOINT

Student Textbook page 157

An individual isotope has an atomic mass, not an average atomic mass. By definition, the atomic mass of carbon-12 is 12u. For other isotopes, the masses of the subatomic particles do not add up to a whole number.

ExpressLab: A Penny for Your Isotopes

Student Textbook page 158
Approximate Time Required: 30 minutes

Tips

- This activity provides students with a simulation of isotopes and average atomic mass using concrete objects they can manipulate. Use Blackline Master 5-4: The Element Centium (Science Inquiry) as a data table to help students organize their findings.
- Each lab group should not have the same ratio of old to new pennies in order for

CHEM FACT

Student Textbook page 159
Boric acid is also used as a fire-retardant in in nickel plating, leather tanning fabrics, and as a fire-retardant in fabrics.

Analysis Question 2 to be valuable.

- About $5 in pennies is required, these can be obtained well ahead of time from class donations or a few rolls can be obtained at the bank.
- Go through the pennies beforehand to sort them into three groups: pre-1982, from 1982 to 1996, and 1997 to the present. Set the first group aside, and make up ten or more bags, depending on the size of the class/lab groups, with different ratios of pre- and post-1997 pennies.
- Try to have the pennies returned, so that you can store them (film containers are good!) for the next time this activity is performed.

Answers to Analysis Questions

1. (a) It is more accurate to get an average mass for each "isotope" since an individual penny could have lost mass (wear) or gained mass (adhering material). Also, the limitations of the balance (a natural thing with any measuring device) produce a larger percentage of uncertainty with a smaller reading. For example, an uncertainty of 0.05 g is a bigger concern with a reading of 1.2 g than with a reading of 12.3 g.
   (b) If you could find the mass of an individual atom, as you can with pennies, step 4 would still be valuable because of the inherent limitations in the accuracy of any measuring device. However, students should realize that all isotopes have exactly the same mass. If they know the mass of one atom of an isotope, they know the masses of any other atom of the same isotope. Meanwhile, the isotopes of centium do not, for reasons mentioned in 1(a).
2. (a) There should be differences since each group had a different isotopic abundance.
   (b) The groups should obtain the same average atomic mass. With normal sized samples, the isotopic abundances are consistent.

Assessment and Evaluation

ThoughtLab/ ExpressLab Investigation	Curriculum Expectations	Assessment Tools / Techniques	Achievement Chart Category	Learning Skills
ExpressLab: A Penny for Your Isotopes, p. 158	*Understanding Basic Concepts* ■ QCR 1.02 explain the relationship between isotopic abundance and relative atomic mass	■ Checklist	■ Knowledge/ Understanding	■ Communication ■ Teamwork

Solutions to Practice Problems

Student Textbook page 160

5.

Isotope	Atomic Mass (u)	Percentage Abundance
H-1	1.0078	x
H-2	2.0140	$100 - x$

Let x be the percentage abundance of H-1. From the Periodic Table, the average atomic mass is aam = 1.01 u and we can write the equation:

$1.01 \text{ u} = [x(1.0078) + (100 \quad x)(2.0140)] \text{ u} / 100$

$101 = 1.0078x + 201.40 - 2.0140x$

$1.0062x = 100.4$

$x = 99.78$

Therefore, natural hydrogen is 99.78% H-1 and 0.022% H-2.

6. Because the Periodic Table gives the average atomic mass of lanthanum as 138.91 u, the same as the atomic mass of La-139 to five significant digits, you know that the abundance of La-138 must be very low.

7.

Isotope	Atomic Mass (u)	Percentage Abundance
Rb-85	84.91	x
R-87	86.91	100 – x

aam = 85.47 u (given in question)

$85.47 \text{ u} = [x(84.91) + (100 \quad x)(86.91)] \text{ u} / 100$

$8457 = 84.91x + 8691 - 86.91x$

$2x = 144$

$x = 72.0$

Therefore, the abundance of Rb-85 is 72%

8.

Isotope	Atomic Mass (u)	Percentage Abundance
O-16	5.9949	x
O-17	16.9991	0.037
O-18	17.9992	100 – (x + 0.037)

aam = 16.00 u (from Periodic Table)

$16.00 \text{ u} = [x(15.9949) + (0.037)(16.9991) + (99.963 - x)(17.9992)] \text{ u} / 100$

$1600 = 15.9949x + 0.6289 + 1799.254 - 17.9992x$

$2.0043x = 199.8829$

so $x = 99.72$

Therefore, natural oxygen is 99.72% O-16, 0.037% O-17 and 0.243% O-18.

Section Review Answers

Student Textbook page 160

1. The average atomic mass is a weighted average of the atomic masses of the isotopes of potassium. Any given atom would have the atomic mass of one of the isotopes, not the average.

2.

Isotope	Atomic Mass (u)	Percentage Abundance
Mg-24	23.985	78.70
Mg-25	24.985	10.13
Mg-26	25.983	11.17

$aam = [78.70(23.985) + 10.13(24.985) + 11.17(25.983)] \text{ u} / 100$

$aam = [1887.6 + 253.1 + 290.2] \text{ u} / 100$

so $aam = 24.31$ u

Therefore, the average atomic mass of magnesium is 24.31 u.

3. No, it is not possible. There would be two unknowns in just one equation. One percentage abundance would have to be known. See question 7 in Practice Problems.
4. Radioisotopes occur in nature along with the stable isotopes of an element. The nucleus of a radioisotope is continually altered due to the loss of alpha or beta particles. This in turn affects the atomic mass of that isotope as it undergoes radioactive decay. Changing masses such as this should not be taken into consideration when calculating the average atomic mass of an element. Another reason these isotopes are not used in calculating the average atomic mass of silver is that their natural abundance is so small, their contribution to the average atomic mass is insignificant.

5.2 The Avogadro Constant and the Mole

Student Textbook pages 161–169

In this section, the mole is defined and the Avogadro constant is introduced. The incredible size of the Avogadro constant is explored. Students then learn how to switch between numbers of moles and numbers of entities using the relationship $n = N / N_A$.

Figure 5.5
Student Textbook page 161

Chemists have to work with large numbers of atoms since even the smallest quantities that can be measured conveniently contain huge numbers of atoms. An indirect measurement is required. We will soon see that the unit chemists use, the mole, is defined in terms of mass in such a way that one mole of any substance contains the same number of particles of that substance.

[Figure to come]

Grouping for Convenience

Student Textbook pages 161–162

Science Background

In science, as in life, special units are developed to make counting objects an easy task. In chemistry, evidence of this way of thinking is found in the mole. The mole is a quantity that represents 6.02×10^{23} particles of an element, molecule, formula unit, or ion. It is used because dealing with individual atoms is very difficult and extremely inconvenient. In most cases, 1 g of any element will contain in the order of 10^{22} atoms, making measuring and counting atoms an insurmountable task. Here are two other examples of how grouping large numbers for convenience is used in science:

- The Astronomical Unit, used to measure the distance between objects in the universe, is equal to the distance between the Earth and the Sun. 1 AU = 150 000 000 km
- The Coulomb, a unit of charge, is used in measuring electrical quantities. Since it is very inconvenient to measure individual electrons moving within a electrical circuit because there are so many of them, scientists measure the number of moving or stored electrons in Coulombs, where 1 C = 6.25×10^{18} electrons.

Teaching Strategies

- Introduce your lesson on the mole by saying to the class, "Yesterday, I had a dozen." Students should quickly point out that the statement needs more information to be useful. Was it a dozen donuts, roses, phone calls from boy/girlfriends? Introduce the new term, mole, as the chemist's dozen.
- It is worth discussing with students that terms like dozen, gross, and mole indicate a number of items, but the type of items should be specified. For example, "a mole of oxygen" is not specific enough. It could refer to oxygen atoms, oxygen molecules, or oxygen ions, so students must include that information.
- Invite students to consider a dozen double-yolk eggs. Ask: How many dozen shells, whites, and yolks are there in a dozen of these eggs? (1, 1, 2.) Does this mean that one dozen therefore equals four dozen? (No, the "things" being counted are not the same.) How many dozen atoms are there in a dozen CO_2 molecules? (3) How many atoms are there? (36) Students can answer these questions easily, and this prepares them for similar questions with numbers of moles.
- Stress that a dozen is an amount, and 12 is the number of things in a dozen.
- Hold a class discussion on other quantities used for groups of things, quantities such as the gross and the ream. Point out that these are used with large scale objects that can be counted easily.

Common Misconceptions

It is a subtle but important point that a dozen is an amount and that 12 is the number of things in a dozen, which is not exactly the same thing. This may seem unimportant with the dozen, but it is important with the mole. It is worth pointing out that, if it were convenient, we could deal with the amount known as the dozen in terms of mass or volume, not in terms of number of things. This is certainly what we will do with the mole.

SUPPORTING DIVERSE STUDENT NEEDS

Many ESL students might have a difficult time understanding the meaning of the term "dozen." If this is the case, bring in a box of a dozen donuts to show the students. Then extend the activity by showing them 12 pictures of common vehicles (make sure each has four wheels – no motorcycles or tractor trailers) and ask them how many tires are on each vehicle? (4 tires.) How many dozen tires are found in all 12 pictures? (4 dozen.) How many tires are there in total among the 12 pictures? (48 tires.)

Language LINK

Student Textbook page 162

The term mole is generally said to be an abbreviation of molecule, but whether mole comes from molecule or directly from the Latin word for mass depends on which dictionary you read. One text—*General Chemistry* (P. W. Atkins (New York: Scientific American Books, 1989))—even claims that the name comes from the Latin word for "massive heap." Discuss with students why the term is an unfortunate choice (it can lead to confusion with the term molecule.) Also, having a three-letter symbol, mol, for a four letter word may seem strange to students.

The Definition of the Mole, The Chemist's Dozen, and How Big is the Avogadro Constant?

Student Textbook pages 162–163

Science Background

The definition of the mole is an amount of a substance that has as many elementary entities as there are atoms in 12 g of carbon-12. One mole of any substance has the same number of entities in it. To understand the need for a chemical unit to measure chemical entities, consider the reaction in which one sodium atom reacts with one chlorine atom to form one formula unit of sodium chloride. The sodium and chlorine atoms react in a 1:1 ratio to form sodium chloride. If one were to carry out this reaction, they would have to measure out 22.99 u of solid sodium and 35.45 u of chlorine gas to produce 58.44 u of sodium chloride. Measurements of such small quantities are not feasible in even the most technologically advanced chemistry labs. Therefore, this measurement ratio is not useful.

By defining a unit, such as the mole, it is easy to combine a large number of sodium atoms with a large number of chlorine atoms. Or, combine one mole of sodium atoms with one mole of chlorine atoms to make 1 mol of sodium chloride formula units.

The number of entities that the mole represents was not known when the concept was defined, but since relative atomic masses were known, mass could be used as an indirect counting system. Sample mass, m, and molar mass, M, can be used to yield number of moles, n: $n = m/M$

This relationship is discussed in Section 5.3. The ratio of numbers of moles of substances in reactions is the same as the ratio of numbers of entities reacting. This was useful even without knowing the Avogadro constant.

Teaching Strategies

- To introduce the mole, write "12 is not enough." on the board. Define elementary chemical entities such as atoms, molecules, and ions. Discuss the need for a big number for such small chemical entities. Define the mole (the amount of a substance containing as many elementary entities as there are atoms in exactly 12 g of carbon-12).
- Introduce the Avogadro constant (6.02×10^{23}). It is worth pointing out that its value was determined after the mole concept was introduced. After all, who would have chosen such a number? Explain that the mole was defined as an amount of matter based on mass—the number of entities in a mole was determined later and named to honour Avogadro. The Avogadro constant was not chosen but determined from experiment. If we were to choose a number to represent the mole, we would most likely choose a number that allows for mathematical ease. These days, the numbers 12 or 144 would not be chosen if we were defining the dozen or gross because it is much easier to work with values based on the number ten.
- The word "entities" can be substituted with the term "chemical unit" as this term clearly describes what the mole represents. A chemical unit can include atoms,

molecules, formula units, or ions.

- Explain to students that the Avogadro constant is so immense that we really cannot imagine its magnitude. The best we can hope for is an appreciation of its vast size. Give students a couple of examples and then have them brainstorm comparisons of their own.
- To have the students begin to appreciate the incredible size of the Avogadro constant, work through, on the board, the solution for the Sample Problem on page 164 of the student textbook.
- Once the large number 6.02×10^{23} has been introduced, it would be a good idea to have students examine and complete the Technology Link on page 165 of the student textbook. This will help students better understand how to use the scientific notation function on their calculator and will be of use in computational problems.
- Return to examples relating the unit dozen to individual units if students are having difficulties with parallel examples involving moles.
- To assist students with the concept of the Avogadro constant and to serve as a reminder, give students Blackline Master 5-5: Chemical Entities and the Avogadro Constant (Information Handout).

SUPPORTING DIVERSE STUDENT NEEDS

For ESL students, write this "graffiti" on the board: "Got mole problems? Call Avogadro 602-1023." Ask students to brainstorm other slogans pertaining to the mole as alternative methods for remembering Avogadro's constant.

For gifted students who have grasped the concept of the mole, an additional project on Avogadro's life and works can be done. A study of Avogadro's work with gases and the relationships between the mole and gas volumes could prove quite interesting for many students.

Common Misconceptions

Students may think the mole was defined with the choice of a number since they are so used to a dozen and because the student textbook mentions the Avogadro constant right before the definition of the mole. It is important for students to realize that when the mole concept was developed, the actual number of entities was unknown. Even now, we cannot count these entities directly, and the Avogadro constant is so immense that a lifetime would not be enough to count that high.

- The mole is the amount of a substance that contains 6.02×10^{23} of the particles represented by its symbol or formula. Confusion can arise from this statement as some substances can exist in more than one form. For example, nitrogen can be found as an atom (N) or as a molecule (N_2). In this case, it is necessary to

Internet LINK

Student Textbook page 163

Be sure to check this site yourself. There are a number of ways that have been used to determine the Avogadro constant. The most direct way is to use electrochemistry, as described briefly in Blackline Master 5-5: Chemical Entities and the Avogadro Constant (Information Handout). A very reliable method for calculating N_A involves the X-ray diffraction of crystals. Other methods, slightly less accurate (giving results accurate within a few percentage points) involve radioactive decay and gas viscosity. Details about the mole and N_A may be found at http://www.moleday.org.

[Figure to come]

Figure 5.7

Student Textbook page163
Each substance shown contains particles different from the particles of the other substances. Each kind of particle has a different mass than the others. Collections of equal numbers of these different particles will therefore have different total masses.

indicate whether the atomic (N) or the molecular (N_2) form is meant.

- Many students will refer to the value of N_A as either the Avogadro constant or the Avogadro number but do not understand that they represent the same quantity. In the student textbook, N_A is correctly referred to as the Avogadro constant, but you may also see N_A referred to as Avogadro's number. There is no difference in these two quantities. The value of N_A does not vary from its assigned value of 6.02 x 10^{23} and for this reason it is referred to as a constant and is found in many tables of scientific constants.

Figure 5.8

Student Textbook page 163

A five loonie stack measures 8.83 mm (using calipers) so a mole stack would measure $[(6.02)(10^{23}) / 5] \times [8.83 \text{ mm}] \times [1 \text{ km} / 105 \text{ mm}]$
$= (1.06)(10^{19})$ km

[Figure to come]

Math LINK

Student Textbook page 164

Using standard calculations for compound interest the amount would be:
$\$(6.02)(10^{23})(1.01^{10}) =$
$\$(6.65)(1023)$

Solutions for Practice Problems

Student Textbook pages 164–165

9. Length of coastline = $(1.7856)(10^4)$ km = $(1.7856)(10^7)$ m
Total length of the sticks = $(6.02)(10^{23})$ m
Number of rows = $(6.02)(10^{23}) / (1.7856)(10^7) = (3.37)(10^{16})$
10. Area of Nunavut = $(1.936113)(10^6)$ km^2 = $(1.936)(10^6)(10^5)^2$ cm^2 = $(1.936)(10^{16})$ cm^2
Area of pastry = $(6.02)(10^{23})(30)^2$ cm^2 = $(5.418)(10^{26})$ cm^2
Number of layers of pastry = $(5.418)(10^{26}) / (1.936)(10^{16}) = (2.8)(10^{10})$
11. distance = (speed)(time) = (100 km / h)[$(6.02)(10^{23})$ days (24 h / day)] = $(1.44)(10^{27})$ km
12. amount = (rate)(time), so time = amount / rate
time = $\$(6.02)(10^{23})$ / [(\$1 / s)(3600 s / h)(24 h / day)(365 day / yr) = $(1.91)(10^{16})$ yr

Converting Moles to Number of Particles, and Converting Number of Particles to Moles

Student Textbook pages 165, 167

Science Background

The Avogadro constant was not chosen, and the mole was defined in terms of relative atomic masses of various elements. Once physicists found the elementary charge and could, in effect, count electrons, the Avogadro constant was determined and the following formulas became valid:

$n = N / N_A$

where

n = number of moles

N = number of entities

N_A = the Avogadro constant.

Although the value N_A is not normally used in the lab, knowing its value is of use for examining chemical phenomena at the atomic level. By understanding the value of N_A, molar quantities can be translated into values that explain observations at the atomic level.

Teaching Strategies

- Write the formula $N = n \times N_A$ on the blackboard. Review each term and the related units with students. Be very clear what units are required for each variable.
  n = number of moles (mol)
  N = number of entities (atoms, molecules, formula units)
  N_A = Avogadro's constant
- Develop a relationship between moles and number of particles in the same way that a student understands that there are 1000 m in 1 km. To convert between m and km, you must either multiply by 1000 or divide by 1000. Refer to this 1000 as a conversion factor. The same train of thought applies to the mole. N_A is just a conversion factor between moles and number of particles.

  meters $\xrightarrow{\times 1000}$ kilometers

  meters $\xrightarrow{\div 1000}$ kilometers

  moles $\xrightarrow{\times N_A}$ number of particles

  moles $\xrightarrow{\div N_A}$ number of particles

- Have students work through the Sample Problems on pages 166 and 168 of the student textbook. Ask students how these problems, as well as their solutions, are similar. Both problems involve the moles, number of particles and Avogadro's constant. One of the sample problems uses the formula in its normally stated form,
  $N = n \times N_A$
  while the other requires a simple re-arrangement to the following form:
  $n = N / N_A$
- The student textbook provides a number of practice problems on converting

SUPPORTING DIVERSE STUDENT NEEDS

Students who are having trouble mastering formula manipulation should be given additional practice in this skill. Blackline Master 5-6: Formula Manipulation (Skill Builder), would be a useful tool for helping these students.

between moles and number of particles. Use these problems to reinforce the simplicity of the conversion between these two quantities. At the same time, stress the importance of using proper problem solving techniques and the inclusion of all units in solving for the final answer.

- Remember that units can be manipulated in a mathematical problem just as numbers are and they can confirm a correct answer or draw attention to an incorrect solution.

Common Misconceptions

- When solving problems with units, students will invariably attempt to multiply a numerical by a unit. This is not possible. Try working through problems involving units just like solving polynomial problems in math class. For each term in a polynomial, there is a numerical co-efficient and a literal co-efficient. In the case of the mole, the Avogadro constant, 6.02×10^{23} is the numerical co-efficient and the unit – particles/mole – is the literal co-efficient. When solving polynomials, students have been taught to work with the numerical co-efficients alone and the literal co-efficients on their own. The answer is then expressed as a combination of both co-efficients. Looking at a problem in this manner might help students in cancelling out units to arrive at the desired quantity.

ThoughtLab: The Magnitude of the Avogadro Constant

Student Textbook page 165
Approximate Time Required: 30 minutes (if all data is provided), 60–75 minutes (if reference material is readily available)

Tips

- This activity provides reinforcement concerning the size of the Avogadro constant, and is therefore very important. However, if you need to minimize class time spent on it, either provide the necessary data or assign groups to collect data outside class time, and report findings to the class.
- The time required for this depends on how you wish to use the lab, taking into account your school's resource centre and Internet availability.
- The data needed with regard to the golf balls, currency, teaspoons, and apples can be obtained fairly quickly in class with standard measuring devices.
- You can provide the geographical and astronomical data, or students can collect it themselves as a homework assignment.
- Note that the values for geographical data are time dependent as new measurement techniques (lasers, satellites) are employed. Students should use the newest data.
- The answers below reflect the data included, so answers will vary, but students' answers should be of the same order of magnitude.

Answers to Analysis Questions

1. Canada's land surface area = $(9.215)(10^6)$ km^2 = $(9.215)(10^6)(10^5)^2$ cm^2
   Golf ball diameter = 4.1 cm. Assume each ball "covers" a 4 cm by 4 cm square or 16 cm^2
   Number of balls to cover land surface once is $(9.215)(10^{16})$ / 16 = $(5.76)(10^{15})$
   Number of layers to use a mole of balls is $(6.02)(10^{23})$ / $(5.76)(10^{15})$ = $(1.05)(10^8)$
   So the coating would be about 100 million balls thick.
2. Mean Earth-Moon distance = $(3.844)(10^8)$ m = $(3.844)(10^{10})$ cm
   Length of $5 bill = 15 cm
   Number of bills to reach moon and back = $2(3.844)(10^{10})$ / 15 = $(5.13)(10^9)$
   Number of round trips with a mole of bills = $(6.02)(10^{23})$ / $(5.13)(10^9)$ = $(1.17)(10^{14})$
3. Volume of oceans = $(1.347)(10^9)$ km^3 = $(1.347)(10^9)(10^5)^3$ cm^3 = $(1.347)(10^{24})$ mL
   Volume of a teaspoon = 5 mL
   Volume of a mole of teaspoons = $(6.02)(10^{23})(5)$ mL = $(3.01)(10^{24})$ mL
   Thus the oceans would be completely drained.
4. Mass of Earth = $(5.977)(10^{24})$ kg
   Mass of typical apple = 200 g = 0.2 kg
   Mass of a mole of apples = $(6.02)(10^{23})(0.2)$ kg = $(1.2)(10^{23})$ kg
   Thus, the mass of Earth is roughly an order of magnitude larger than the mass of a mole of apples.
5. Current world population (January 2001) = $(6.13)(10^9)$
   Number of Earth-sized planets needed for a mole of humans = $(6.02)(10^{23})$ / $(6.13)(10^9)$ = $(9.8)(10^{13})$
   Roughly 10^{14} planets would be needed.

Assessment and Evaluation

ThoughtLab/ ExpressLab Investigation	Curriculum Expectations	Assessment Tools / Techniques	Achievement Chart Category	Learning Skills
ThoughtLab: The Magnitude of the Avogadro Constant, p. 165	*Overall Expectations* ■ QCR V.01 demonstrate an understanding of the mole concept and its significance in the analysis of chemical systems *Understanding Basic Concepts* ■ QCR 1.01 demonstrate an understanding of Avogadro's number, the mole concept, and the relationship between the mole and molar mass *Developing Skills of Inquiry and Communication* ■ QCR 2.03 solve problems involving quantity in moles, number of particles, and mass.	■ Checklist	■ Knowledge/ Understanding ■ Communication	■ Teamwork ■ Work Habits

Language LINK

Student Textbook page 166

A billion is different in the United States—a thousand million or 10^9—and the UK—a million million or 10^{12}—so specify the meaning for the students. The American version is more common in Canada, so the Avogadro constant is 23-9 or 14 orders of magnitude greater.

History LINK

Student Textbook page 168

Avogadro's hypothesis was a key to allowing the development of the mole concept. It was later formalized into a mathematical equation called Avogadro's Law. When it is covered in Unit 4, this connection can be explored.

Solutions for Practice Problems

Student Textbook pages 166–167

These solutions use the formula:
$N = (n)(N_A)$
where
N is the number of entities
n is the number of moles
N_A is the Avogadro number
Note: All solutions begin with a list of information from the question.

13. $n = 0.0178$ mol; $N = ?$; $N_A = (6.02)(10^{23})$ atoms/mol
$N = (n)(N_A) = (0.0178)(6.02)(10^{23})$ atoms $= (1.07)(10^{22})$ atoms
There are $(1.07)(10^{22})$ iron atoms in the pin.
14. $n = (4.70)(10^{-4})$ mol; $N = ?$; $N_A = (6.02)(10^{23})$ atoms/mol
$N = (n)(N_A) = (4.70)(10^{-4})(6.02)(10^{23})$ atoms $= (2.83)(10^{20})$ atoms
There are $(2.83)(10^{20})$ gold atoms in the sample.
15. $n = 0.21$ mole; $N = ?$; $N_A = (6.02)(10^{23})$ formula units/mol
$N = (n)(N_A) = (0.21)(6.02)(10^{23}) = (1.3)(10^{23})$ formula units
There are $(1.3)(10^{23})$ formula units in the sample.
16. $n = 55.6$ mol; $N = ?$; $_A = (6.02)(10^{23})$ molecules/mol
$N = (55.6)(6.02)(10^{23})$ molecules $= (3.35)(10^{25})$ molecules
There are $(3.35)(10^{25})$ water molecules in one litre of water.
17. Ethyl acetate, $C_4H_8O_2$
(a) 2.5 mol; $= ?$; $N_A = (6.02)(10^{23})$ molecules/mol
$N = (2.5)(6.02)(10^{23})$ molecules $= (1.5)(10^{24})$ molecules
There are $(1.5)(10^{24})$ molecules in the bottle.
(b) There are 14 atoms per molecule, so the number of atoms in the bottle is $(14)(1.5)(10^{24})$ or $(2.1)(10^{25})$.
(c) There are four C atoms per molecule, so the number of C atoms in the bottle is $(4)(1.5)(10^{24})$ or $(6.0)(10^{24})$.
18. (a) Sodium sulfate, Na_2SO_4.
$n = 0.829$ mol; $N = ?$; $N_A = (6.02)(10^{23})$ formula units/mol
$N = (0.829)(6.02)(10^{23})$ formula units $= (4.99)(10^{23})$ formula units
There are $(4.99)(10^{23})$ formula units in the sample.
(b) There are two sodium ions per formula unit, so there are $(2)(4.99)(10^{23})$ or $(9.98)(10^{23})$ sodium ions.

Solutions for Practice Problems

Student Textbook page 168

19. $N = (7.71)(10^{24})$ molecules; $n = ?$; $N_A = (6.02)(10^{23})$ molecules/mol
$n = N / N_A = (7.71)(10^{24}) / (6.02)(10^{23})$ mol $= 12.8$ mol
There are 12.8 moles of aluminum oxide in the sample.
20. $N = (8.03)(10^{26})$ molecules; $n = ?$; $N_A = (6.02)(10^{23})$ molecules/mol
$n = N / N_A = (8.03)(10^{26}) / (6.02)(10^{23})$ mol $= (1.33)(10^3)$ mol
There are $(1.33)(10^3)$ moles of ammonia in the vat.
21. There are three atoms per molecule of HCN so the number of molecules is

$N = (3.33)(10^{22}) / 3 = (1.11)(10^{22})$
Now n = ? and $N_A = (6.02)(10^{23})$
$n = N / N_A = (1.11)(10^{22}) / (6.02)(10^{23}) = (1.84)(10^{-2})$
There are $(1.84)(10^{-2})$ moles of cyanic acid in the sample.

22. In acetic acid there are two carbon atoms per molecule so the number of molecules is
$N = (1.40)(10^{23}) / 2 = (7.0)(10^{22})$ molecules
Now n = ? and $N_A = (6.02)(10^{23})$
$n = N / N_A = (7.0)(10^{22}) / (6.02)(10^{23}) = (1.16)(10^{-1})$ mol
There are $(1.16)(10^{-1})$ moles of acetic acid in the sample.

Section Review Answers

Student Textbook page 169

1. A mole is an amount of matter that always has the same number of entities in it as exactly 12 g of carbon-12. Here are three examples.

one mole of O_2:	one mole of Ag:	one mole of H_2O:
$(6.02)(10^{23})$ molecules	$(6.02)(10^{23})$ atoms	$(6.02)(10^{23})$ molecules

2. Assuming a billion to be (10^9), each person would get $\$(6.02)(10^{23}) / (6)(10^9) = \$(10^{14})$
3. Time for a mole of heartbeats would be
$(6.02)(10^{23})$beats / [60 beats / min] = $(1.0)(10^{22})$min
Converting to years, this is
$(1.0)(10^{22})$ min (1h / 60min)(1day / 24h)(1yr / 365day) = $(1.9)(10^{16})$ years.
4. N = ?; n = 3.45 mol; $N_A = (6.02)(10^{23})$atoms/mol
$N = (n)(N_A) = (3.45)(6.02)(10^{23})$ atoms = $(2.08)(10^{24})$ atoms
5. (a) $N = (2.56)(10^{24})$ molecules; $N_A = (6.02)(10^{23})$ molecules/mol; n = ?
$n = N / N_A = (2.56)(10^{24})$ mol / $(6.02)(10^{23})$ = 4.25 mol
There are 4.25 moles of carbon dioxide.
(b) There are three atoms per molecule, so there are (3)(4.25) or 12.8 moles of atoms present.
6. n = 0.5 mol; $N_A = (6.02)(10^{23})$ molecules / mol; N = ?
$N = (n)(N_A) = (0.5)(6.02)(10^{23})$ molecules = $(3.0)(10^{23})$ molecules
Since helium molecules are monatomic, there are $(3.0)(10^{23})$ atoms of helium in the balloon.
7. (a) n = 5.69 mol; $N_A = (6.02)(10^{23})$ molecules/mol; N = ?
$N = (n)(N_A) = (5.69)(6.02)(10^{23})$ molecules = $(3.43)(10^{24})$ molecules
The number of molecules is $(3.43)(10^{24})$.
(b) Since there are six hydrogen atoms per benzene molecule, the number of hydrogen atoms is $(6)(3.43)(10^{24})$ or $(2.06)(10^{25})$.
8. (a) n = 1.17 mol of Al_2O_3; N = ?; $N_A = (6.02)(10^{23})$ molecules/mol
$N = (n)(N_A) = (1.17)(6.02)(10^{23})$ molecules = $(7.04)(10^{23})$ molecules
There are $(7.04)(10^{23})$ molecules in the sample.
(b) Each molecule contains five atoms, so the number of atoms is $(5)(7.04)(10^{23})$ or $(3.52)(10^{24})$.
(c) Each molecule contains three oxygen atoms, so the number of oxygen atoms in the sample is $(3)(7.04)(10^{23})$ or $(2.11)(10^{24})$.

Internet LINK

Student Textbook page 168

Here is a table of the SI base units from which all other units are derived.

Physical Quantity	Base Unit
Length	metre (m)
Mass	kilogram (kg)
Time	second (s)
Electric Current	ampere (a)
Temperature	kelvin (K)
Amount of Substance	mole (mol)
Luminous Intensity	candela (cd)

More details are available in *Metric Style Guide* (Council of Ministers of Education (Canada), Toronto, 1975). This was distributed to schools across Canada.
Information on SI base units can be found on the Internet at:
http://physics.nist.gov/cuu/Units/index.html.

9. They defined it this way for two reasons. First, they wanted it to contain the same number of entities all the time just as a dozen does. Secondly they wanted to use the relative mass values for atoms but with quantities that could be measured conveniently in the lab. Thus they chose 12 g of carbon-12 as the basis of the definition. Students may also say that scientists used an isotope of carbon as a point of reference because it is so common and because it is the building block for all organic matter.
10. Number of moles of ZnO = $(3.28)(10^{24}) / (6.02)(10^{23}) = 5.45$ moles.
    The metal sample contains only 2.78 moles of zinc atoms while the sample of zinc oxide contains 5.45 moles of ZnO. Since one molecule of ZnO contains one atom of Zn, then if there are 5.45 moles of ZnO present there must be 5.45 moles of Zn in sample. Therefore the compound contains more zinc.

5.3 Molar Mass

Student Textbook pages 170–182

In section 5.2, the relationship $n = N / N_A$ was explored. In this section, molar mass is defined, and the relationship $n = m / M$ is developed. Then both relationships are used to link number of entities, number of moles, and mass. After molar mass is introduced, this section discusses variants of essentially the same problem—namely the conversion from one description of an amount of matter to an equivalent description of the same amount of matter.

Mass and the Mole; What is Molar Mass?; Finding Molar Mass of Compounds

Student Textbook pages 170–171

Science Background

Atomic mass is the mass of one atom measured in u, as recorded in the Periodic Table. Molecular mass is the mass of one molecule measured in u. Formula mass is the mass of one formula unit measured in u. The second and third of these masses are found by adding the relevant atomic masses.

The use of measurements in atomic mass units is not feasible within a laboratory environment. To extend the mole to a scale suitable for measurement, the element carbon-12 must be examined further. Each atom of C-12 is known to have a mass of 12 u or 1.99×10^{-23} g. If the mass of one mole of carbon-12 atoms were determined, the total mass of carbon atoms would be equal to
$(1.99 \times 10^{-23}$ g/atom$)(6.02 \times 10^{23}$ atoms/mole$) = 12$ g/mole

So if the mass of one carbon atom is equal to 12 u, and the mass of one mole of carbon atoms is equal to 12 g, then it can be stated that: the mass of one mole of any substance (expressed in grams) is numerically equal to the atomic mass of that same substance expressed in amu. Students may also say that scientists used an isotope of carbon as a point of reference because it is so common and because it is the building block for all organic matter. Similarly, the molar mass represents the mass per mole of a substance (in units of grams per mole). That is, one mole of C-12 = 12 g or the molar mass of C-12 is 12 g/mol.

Earlier names for molar mass include gram atomic weight, gram molecular weight, and gram formula weight. These have been replaced by the umbrella term, the mole. As the old terms imply, the mole is a mass, measured in grams and using the numerical values of the average atomic masses discussed in the previous section.

Teaching Strategies

- The molar mass of any substance can be calculated in the same manner as the average atomic mass of the same substance. Explain to the students that the average mass of a particle expressed in atomic mass units can be used to express the mass of one mole of that same substance in grams. Remind students that the average atomic masses are used because large samples contain different isotopes of any given element.
- The atomic mass on the Periodic Table represents the average mass of an atom in an element. It also represents the mass of one mole of atoms of that element, expressed in grams. After explaining this concept to students, it might be helpful to place a number of samples around the room to show exactly how much space one mole of a particular element takes up (e.g 32.07 g of S, 55.85 g of Fe).
- The definition of molar mass allows for the calculation of the size of an atomic mass unit (u) in grams. To demonstrate this concept to students, ask them to consider any element X with atomic mass x u, molar mass x g. The following equation is trivial but useful. Mass of one mole = mass of one mole. That is: x g = $(6.02)(10^{23})(x)$ u. We see that x can be divided out, making the choice of element unimportant. We obtain 1 g = $(6.02)(10^{23})$ u, which means that 1 u = $(1.66)(10^{-24})$ g.
- Knowing the molar mass of any element or particle is a very handy tool for chemists as it provides them with a conversion factor to move between mass and moles, in the same manner that N_A is a conversion factor between moles and the number of particles of a particular substance.
- Blackline Master 5-7: Molar Mass Computations (Overhead Master), can be used to show students the relationships between atomic mass and molar mass for elements and compounds.
- Blackline Master 5-8: Molar Mass (Problem Solving), provides further practice in determining the molar mass for various elements/compounds.

Common Misconceptions

Some students have trouble seeing relative masses being useful for counting and think that they need to know the number amount. To deal with this, have them consider the building site for a new house. Each concrete block (for the basement) has three times the mass of each brick (for the outer skin of the house). It turns out, for this house, that equal numbers of bricks and blocks are needed. If the suppliers deliver a 1000 kg pile of bricks, what mass of blocks must they deliver to supply the same number of blocks? (3000 kg) Point out that they did not need to know the number of bricks (or blocks) to answer this.

SUPPORTING DIVERSE STUDENT NEEDS

ESL students may not be able to understand what is meant by individual chemical entities. These students may benefit from samples of molar amounts of various elements/compounds placed throughout the classroom (labeled with their molar masses). Have students move around the room in pairs and describe to one another what is represented in each pile (e.g. type of element/compound, types of atoms present, molar mass of element/compound, and number of particles present).

Investigation 5-A:

Modelling Mole and Mass Relationships

Student Textbook pages 172–173
Approximate Time Required: 60 minutes

Tips

- Introduce this investigation by pointing out that it shows us how to count everyday objects by indirect means. This is not something we need to do very often, but this ability is essential for counting chemical entities. Have the students carry out the investigation, recording their data on Blackline Master 5-9: Indirect Counting (Science Inquiry). This data table can form a portion of their investigation reports, which should be collected for assessment.
- Stress the fact that opening the film canisters takes away the whole point of the exercise.
- Be sure the rice does not end up on the floor (safety) or in the sinks (clogged drains).
- Use small nuts and washers and put a student in charge of distributing and collecting them to reduce loss.
- If you do not want to deal with nuts and washers, try two types of small candy, this way clean-up will not be too difficult. Remember that students should never eat anything that they have been working with in the laboratory.

Answers to Analysis Questions

Student Textbook pages 172–173

1. The mass is small and the rice grains are not identical. Using a larger sample gets a more accurate mass for the sample (balance accuracy) and a more accurate average mass per grain.
2. Knowing the average mass of a rice grain and the total mass of the sample allows the calculation: number of grains = total mass / mass of one grain.
3. Divide $(6.5)(10^3)$ g by your value of the mass of one grain to find how many grains there would be.
4. (a) number of atoms = $(23.8 / 4.00)(6.02)(10^{23}) = (3.58)(10^{24})$.
   (b) $n = N / N_A$ and $n = m / M$.
   (c). See 2 above.
5. (a) You cannot find the number of washers without more data.
   (b) You would need to know the actual mass of either a nut or a washer and the mass of the empty container.
6. (a) Carbon Oxygen $12 / 5.8 = 32.0 / x$
   12.0 g 32.0 g $x = 15.5$
   5.8 g x g
   Therefore, 15.5 grams of oxygen will react.
   You know that masses in the ratio 12 /32 will have equal numbers of entities.
   (b) In both cases, you use relative masses for indirect counting to obtain equal numbers of things, be they nuts and washers, or atoms and molecules.

Answer to Conclusion Question

7. Answers should reflect the idea that the number of moles can be linked to both number of entities and mass, in effect, to the two formulas in 4(b) above or to the list called the "mole box" in Teaching Strategies.

Answers to Applications Questions

8. Let students know that this question refers to Analysis Question 6(a). They should not use the Avogadro constant in their calculations, instead they use indirect counting and proportional reasoning.
9. Most often, the relationship of mass to number of moles is used.

Assessment and Evaluation

ThoughtLab/ ExpressLab Investigation	Curriculum Expectations	Assessment Tools / Techniques	Achievement Chart Category	Learning Skills
5-A: Modelling Mole and Mass Relationships, pp. 172–173	*Understanding Basic Concepts* ■ QCR 1.01 demonstrate an understanding of Avogadro's number, the mole concept, and the relationship between the mole and molar mass ■ QCR 1.05 state the quantitative relationships expressed in a chemical equation (e.g., in moles, grams, atoms, ions, or molecules) *Developing Skills of Inquiry and Communication* ■ QCR 2.03 solve problems involving quantity in moles, number of particles, and mass	■ Rubric	■ Inquiry ■ Communication	■ Organization ■ Work Habits

Solutions for Practice Problems

Student Textbook page 174

23. Reading from the Periodic Table
    (a) Xe: M = 131.29 g/mol.
    (b) Os: M = 190.23 g/mol.
    (c) Ba: M = 137.33 g/mol.
    (d) Te: M = 127.60 g/mol.
24. Adding values from the Periodic Table
    (a) NH_3: $M = 14.01 + 3(1.01) = 17.04$ g/mol.
    (b) $C_6H_{12}O_6$: $M = 6(12.01) + 12(1.01) + 6(16.00) = 180.2$ g/mol.
    (c) $K_2Cr_2O_7$: $M = 2(39.10) + 2(52.00) + 7(16.00) = 294.2$ g/mol.
    (d) $Fe_2(SO_4)_3$: $M = 2(55.85) + 3[32.07 + 4(16.00)] = 399.91$ g/mol.

25. $SrSO_4$: M = 87.62 + 32.07 + 4(16.00) = 183.69 g/mol.
26. $[Cu(NH_3)_4]^{2+}$: M = 63.55 + 4[14.01 + 3(1.01)] = 131.71 g/mol.

Counting Particles Using Mass, Converting from Moles to Mass, and Converting from Mass to Moles

Student Textbook pages 174, 175–176

Teaching Strategies

CHECKPOINT

Student Textbook page 175

$CaCl_{2\,(aq)} + (NH_4)_2SO_{4\,(aq)} \rightarrow CaSO_{4\,(s)} + 2NH_4Cl_{(aq)}$

This is a double displacement reaction.

- Have students recall and manipulate the formula $n = N/N_A$ again to solve for N and N_A.
- Take a look at the units of molar mass: grams per mole (g/mol). What quantities will these units allow a conversion between? Have students recall that the conversion factor of 2.25 cm/inch allows the conversion between distances in centimeters and inches. Therefore, the unit g/mol should allow the conversion between mass and moles, where g/mol is the molar mass of a given entity.
- If molar mass = M, then $M = \frac{m_{(g)}}{n_{(mo)}}$, by dividing a mass in grams by its amount in moles, a solution is obtained in units of g/mol, or the molar mass.
- Have students examine how the formula $M = \frac{m}{n}$ can be manipulated to solve for n and m.

  $n = m/M$
  $m = n \times M$
- Work through the Sample Problems on pages 175 and 177 with students. This will allow them to become comfortable with converting between moles and mass and vice versa.

Solutions for Practice Problems

Student Textbook page 176

In each case find the molar mass, M, then use $n = \frac{m}{n}$ to solve for m [$m = (n)(M)$]

27. (a) C: M = 12.01 g/mol; n = 3.90 mol; m = (3.90)(12.01) g = 46.8 g
    (b) O_3: M = 3(16.00) = 48.00 g/mol; n = 2.50 mol; m = (2.50)(48.00) g = 120 g
    (c) C_3H_8O: M = 3(12.01) + 8(1.01) + 16.00 = 60.11 g/mol; $n = (1.75)(10^7)$ mol.; $m = (1.75)(10^7)(60.11)$ g = $(1.05)(10^9)$ g
    (d) $(NH_4)_2Cr_2O_7$: M = 2[14.01 + 4(1.01)] + 2(52.00) + 7(16.00) = 252.10 g/mol; $n = (1.45)(10^{-5})$ mol; so $m = (1.45)(10^{-5})(252.10)$ g = $(3.66)(10^{-3})$ g
28. Find the sample masses as in solution 27 and compare the values.
    (a) C: M = 12.01 g/mol; n = 5.00 mol; m = (5.00)(12.01) g = 60.05 g
    Cl_2: M = 70.90 g/mol; n = 1.50 mol; m = (1.50)(70.90) g = 106.35 g
    $C_6H_{12}O_6$: M = 180.18 g/mol; n = 0.50 mol; m = (0.50)(180.18) g = 90.09 g
    The sample of Cl_2 has the largest mass.
    (b) O_2: M = 32.00 g/mol; n = 7.31 mol; m = (7.31)(32.00) g = 234 g
    CH_3OH: M = 32.05 g/mol; n = 5.64 mol; m = (5.64)(32.05) g = 181 g

H_2O: M = 18.02 g/mol; n = 12.1 mol; m = (12.1)(18.02) g = 218 g
The sample of O_2 has the largest mass.

29. Since the density of water is 1.0 g/mL, the mass of 1000 mL is 1000 g or 1 kg.
Using the mole concept:
n = 55.6; M = 18.02 g/mol; m = (55.6)(18.02) g = 1001.9 g

30. n = 255 mol and for C_8H_8: M = 8(12.01) + 8(1.01) = 104.16 g/mol; m = (255)(104.16) g = 26561 g = 26.56 kg
The engineer needs 26.6 kg.

[Figure to come]

Figure 5.13

Student Textbook page 176

Triangles for solving for scientific relationships, similar to that shown in Figure 5.13, can be used in working with current, voltage and resistance ($V = IR$) and density ($D = m / V$).

Solutions for Practice Problems

Student Textbook page 177

In each case, calculate molar mass (M) and use it with sample mass (m) to calculate number of moles (n). Use $n = m / M$

31. (a) Mo: M = 95.94 g/mol; m = 103 g; n = 103 / 95.94 = 1.07 mol
(b) Pd: M = 106.42 g/mol; m = $(1.32)(10^4)$ g; n = $(1.32)(10^4)$ / 106.42 = 124 mol
(c) Cr: M =52.00 g/mol; m = 736 g; n = 736 / 52.00 = 14.2 mol
(d) Ge: M = 72.61 g/mol; m = 0.0563 g; n = 0.0563/72.61 = $(7.75)(10^{-4})$ mol

32. (a) SiO_2: M = 60.09 g/mol; m = 39.2 g; n = 39.2 / 60.09 = 0.652 mol
(b) HNO_2: M = 47.02 g/mol; m = 7.34 g; n = 7.34 / 47.02 = 0.156 mol
(c) CF_4: M = 88.01 g/mol; m = $1.550(10^5)(10^3)$ g; n = 1.550 $(10^5)(10^3)$ / 88.01 = $17.6(10^6)$ mol
(d) C_8H_9I: M = 232.07 g/mol; m = $(8.11)(10^{-3})$g; n = $(8.11)(10^{-3})$ / 232.07 = $(3.49)(10^{-5})$ mol.

33. NaCl: M = 58.44 g/mol; m = 10 000 g; n = 10 000 / 58.44 = $1.7(10^2)$ mol.

34. C_8H_{18}: M = 114.26 g/mol; m = 20 000 g; n = 20 000 / 114.26 = 175 mol.

Chemistry Bulletin

Student Textbook page 178

Answers to Making Connections Questions

1. Vitamins and minerals ingested daily should either be used by the body or eliminated. Build up can be harmful. For example vitamin A is toxic in larger amounts. Polar bear liver is so rich in this vitamin that the Innu do not eat it (it would be lethal). Any vitamins that are fat-soluble are not eliminated by the kidneys and excess amounts accumulate in fatty tissues like the liver.
2. Given that one tablet contains 950 mg of calcium citrate or 0.950 g
Also given that 1.00 g of calcium citrate contains $(5.26)(10^{-3})$ moles of calcium.
Thus each tablet contains $(0.950/1.00)(5.26)(10^{-3})$ moles or $(4.997)(10^{-3})$ moles of calcium.
The molar mass of calcium is 40.08 g/mol
Using $m = \dfrac{n}{M}$, each tablet contains $(4.997)(10^{-3})(40.08)$g or 0.200g or 200mg.
In order to meet the daily RNI, 4 or 5 tablets should be taken (800 or 1000 mg).

Converting Between Moles, Mass, and Number of Particles

Student Textbook page 179

Teaching Strategies

Math LINK

Student Textbook page 179

Use two decimal places of atomic mass values. This will give four or five significant digits except for elements one to four where it gives three significant digits. At the end of a calculation, round off the number of significant digits to match those provided in the problem.

- Now that students have been introduced to the relationships between moles and numbers [$n = N / N_A$] and between moles and mass [$n = m / M$], they are ready to link all three of these things together. The Concept Organizer feature on page 182 of the student textbook is one approach to this.
- Another approach is provided by Blackline Master 5-10: Mass, Moles and Molecules (Overhead Master). It points out that *n* is common to both formulas, which allows for conversions between numbers of entities, moles, and mass. To illustrate how to use this, work through the solution to the Sample Problem on pages 179 and 180.
- Based on the mathematical strength of your students, you may wish to use an alternative approach to the above. For example, we can list equivalent descriptions of one mole of any substance using the "Mole Box."

Mole Box
1 mole
$(6.02)(10^{23})$ entities
molar mass

- An example of the mole box is shown below.

1 mole Ag	1 mole CO_2 (g)
$(6.02)(10^{23})$ atoms	$(6.02)(10^{23})$ molecules
107.87 g	44.01 g

- For additional practice of this, use Blackline Master 5-11: The Mole Box (Skill Builder).
- Provide students with this "rule structure" or algorithm to use once they have read and recognized questions required conversions between moles, mass, and number of particles.
  – Make a data table with two rows labelled "sample" and "mole."
  – On the "sample" row list, with units, the quantity given and the unknown quantity sought.
  – Make a "Mole Box" off to the side, for the substance in question
  – On the "mole" row fill in data from the "Mole Box" so that the units in the columns match.
  – Use ratio and proportion to find the unknown.

  Here is an example:

  There are 28 g of gold plated on a medal given to an athlete. How many atoms of gold are on the medal?

  Solution:

  Sample of gold = 28 g = *x* atoms

  1 mole of gold = 196.97 g = $(6.02)(10^{23})$ atoms

1 mole Au
$(6.02)(10^{23})$ atoms
196.97 g/mol

28 / 196.97 = x / (6.02)(10^{23})
x = (8.56)(10^{22})
There are (8.6)(10^{22}) gold atoms on the medal.

- Work some examples (choosing text questions) with the students until they are comfortable with a problem-solving framework that works for them. Including units in the data table is crucial. The above framework for problem solving leads into one for approaching stoichiometric problems in Chapter 7.

Solutions for Practice Problems

Student Textbook page 180

35. (a) Sample = (6.02)(10^{24}) formula units = x g
Mole = (6.02)(10^{23}) formula units = 136.29 g
(6.02)(10^{24}) / (6.02)(10^{23}) = x / 136.29
x = 1363
The mass is 136(10^{3}) g.

1 mole $ZnCl_2$
(6.02)(10^{23}) formula units
136.29 g

(b) Sample = (7.38)(10^{21}) formula units = x g
Mole = (6.02)(10^{23}) formula units = 811.54 g
(7.38)(10^{21}) / (6.02)(10^{23}) = x / 811.54
x = 9.95
The mass is 9.95 g.

1 mole $Pb_3(PO_4)_2$
(6.02)(10^{23}) formula units
811.54 g

(c) Sample = (9.11)(10^{23}) molecules = x g
Mole = (6.02)(10^{23}) molecules = 483.39 g
(9.11)(10^{23}) / (6.02)(10^{23}) = x / 483.39
x = 731.5
The mass is 732 g.

1 mole $C1_5H_{21}N_3O_{15}$
(6.02)(10^{23}) molecules
483.39 g

(d) Sample = (1.20)(10^{29}) molecules = x g
Mole = (6.02)(10^{23}) molecules = 108.02 g
(1.20)(10^{29}) / (6.02)(10^{23}) = x / 108.02
x = (2.15)(10^{7})
The mass is (2.15)(10^{7}) g or (2.15)(10^{4}) kg.

1 mole N_2O_5
(6.02)(10^{23}) molecules
108.02 g

36. Sample 254 formula units = x g
Mole = (6.02)(10^{23}) = 42.39 g
254 / (6.02)(10^{23}) = 4.22(10^{-22}) mol
4.22(10^{22}) x molar mass of lithium =
4.22(10^{-22}) mol x 6.94 g/mol = 2.93(10^{-21}) g.

1 mole LiCl
(6.02)(10^{23}) formula units
42.39 g

37. Sample = 1 atom = x g
Mole = (6.02)(10^{23}) atoms = 47.87 g
1 / (6.02)(10^{23}) = x / 47.87
x = (7.95)(10^{-23})
The mass is (7.95)(10^{-23}) g.

38. Sample = 1 molecule = x g
Mole = (6.02)(10^{23}) molecules = 376.41 g
1/ (6.02)(10^{23}) = x / 376.41
x = (6.25)(10^{-22})
The mass is (6.25)(10^{-22}) g.

1 mole riboflavin
(6.02)(10^{23}) molecules
376.41 g

Solutions for Practice Problems

Student Textbook page 181

39. (a) Sample = 10.0 g = x molecules
Mole = 18.02 g = $(6.02)(10^{23})$ molecules
10.0 / 18.02 = x / $(6.02)(10^{23})$
$x = (3.34)(10^{23})$
There are $(3.34)(10^{23})$ molecules.

1 mole H_2O
$(6.02)(10^{23})$ molecules
18.02 g

(b) Sample = 52.4 g = x molecules
Mole = 32.05 g = $(6.02)(10^{23})$ molecules
52.4 /32.05 = x / $(6.02)(10^{23})$
$x = (9.84)(10^{23})$
There are $(9.84)(10^{23})$ molecules.

1 mole CH_3OH
$(6.02)(10^{23})$ molecules
32.05 g

(c) Sample = 23.5 g = x molecules
Mole = 135.04 g = $(6.02)(10^{23})$ molecules
23.5 / 135.04 = x / $(6.02)(10^{23})$
$x = (1.05)(10^{23})$
There are $(1.05)(10^{23})$ molecules.

1 mole S_2Cl_2
$(6.02)(10^{23})$ molecules
135.04 g

(d) Sample = 0.337 g = x formula units
Mole = 811.54 g = $(6.02)(10^{23})$ formula units
0.337 / 811.54 = x / $(6.02)(10^{23})$
$x = (2.50)(10^{20})$
There are $(2.50)(10^{20})$ formula units.

1 mole $Pb_3(PO_4)_2$
$(6.02)(10^{23})$ formula units
811.54 g

40. Each molecule of sodium glutamate has eight H atoms so there are $8(5.3)(10^4)$ or $(4.2)(10^5)$ atoms.

41. Sample = $6.43(10^{-2})$ g = x molecules
Mole = 283.88 g = $(6.02)(10^{23})$ molecules
0.0643 / 283.88 = x / $(6.02)(10^{23})$
$x = (1.36)(10^{20})$
There are $(1.36)(10^{20})$ molecules.

1 mole P_4O_{10}
$(6.02)(10^{23})$ molecules
283.88 g

42. (a) Sample = $(4.35)(10^{-2})$ g = x formula units
Mole = 122.55 g = $(6.02)(10^{23})$ formula units
$(4.35)(10^{-2})$ / 122.55 = x / $(6.02)(10^{23})$
$x = (2.14)(10^{20})$
There are $(2.14)(10^{20})$ formula units.

1 mole $KClO_3$
$(6.02)(10^{23})$ formula units
122.55 g

(b) Each formula unit of $KClO_3$ consists of two ions, so there are $2(2.14)(10^{20})$ or $(4.28)(10^{20})$ ions in the sample.

Section Review Answers

Student Textbook page 182

1. **Atomic Mass (u)** $\xrightarrow{\text{multiply by } 6.02 \times 10^{23}}$ **Molar Mass (g)**

1 atom

6.02×10^{23} atoms
or 1 mole of atoms

2. NH_3: M = 14.01 + 3(1.01) = 17.04 g/mol
   sample m = 78.6g
   (a) $n = m / M$ = 78.6/17.04 = 4.61 mol.
   (b) Using $n = N / N_A$ find $N = (n)(N_A)$ = (4.61)(6.02)(10^{23}) = (2.78)(10^{24}) molecules.
3. 1 u = 1.66(10^{-24}) g or 1 g = (6.02)(10^{23}) u.
   (a) 28.09 g/mol / 6.02(10^{23})(mol^{-1}) = 4.67 x 10^{-23} g.
   (b) For Si, M = 28.09g = 28.09(6.02)(10^{23}) u = 1.69(10^{25}) u.
4. NaCl: M = 22.99 + 35.45 = 58.44 g/mol
   Sample n = 0.789 mol
   (a) Using $n = m / M$ find $m = (n)(M)$ = 0.789(58.44)g = 46.1 g.
   (b) Using $n = N / N_A$ find $N = (n)(N_A)$ = 0.789(6.02)(10^{23}) = 4.75(10^{23}) formula units.
   (c) There are two atoms per formula unit so the total number of atoms is 2(4.75)(10^{23}) or 9.50(10^{23}).
5. Sample 1 g x atoms
   Mole 12.01 g (6.02)(10^{23}) atoms
   1.00 / 12.01 = x / (6.02)(10^{23})
   x = 5.01(10^{22})
   There are 5.01(10^{22}) carbon atoms in the diamond.

1 mole C (6.02)(10^{23}) atoms 12.01 g

6. A bottle containing 100 tablets contains 200 mg of copper.
   Number of moles of Cu in the 100 tablets
   Molar mass of Cu = 63.55 g/mol
   0.200 g / 63.55 g/mol = 3.15 x 10^{-3} mol of Cu
   If there is 3.15 x 10^{-3} mol of Cu in the bottle, then there must be the same number of moles of CuO present.
   Molar mass of CuO = 79.56 g/mol
   Mass of CuO in the bottle = (3.15 x 10^{-3} mol)(79.56 g/mol) = 0.2506 g or 250.6 mg of CuO.
   Mass of CuO in one tablet = 250.6 mg / 100 tablets = 2.5 mg CuO per tablet.

Chapter 5 Review Answers

Student Book pages 183–185

Reviewing Key Terms

- average atomic mass: weighted average of the mass of the element's isotopes
- Avogadro constant: 6.02 x $10^{23,}$ which is equal to the number of particles found in one mole of any substance
- isotopic abundance: relative amount of each isotope found in an element expressed as a percentage or as a decimal fraction
- mass spectrometer: a powerful instrument that generates a magnetic field to obtain data about the mass and abundance of isotopes; can also find the mass of molecules
- mole: the amount of a substance that contains as many elementary entities (atoms, molecules, or formula units) as exactly 12 g of carbon-12
- molar mass: the mass of one mole of a substance
- weighted average: a statistical calculation taking into account not only the values associated with a set of data but also the abundance or importance of each value

Answers to Knowledge/Understanding Questions

1. Atomic mass is the mass of one atom of a specific isotope. Average atomic mass is the weighted average of atomic masses for the naturally occurring mix of isotopes of an element. For example, Cl-35 has atomic mass 34.97 u while the average atomic mass for chlorine is 35.45u (since there is also Cl-37 to be considered).
2. The average has to be weighted to reflect the relative abundances of the various isotopes, which are not present in equal numbers.
3. Chlorine has two isotopes, Cl-35 and Cl-37. The average atomic mass of chlorine is due to the weighted average of the two isotopes, therefore, no one chlorine atom has a mass of 35.45 u.
4. The Avogadro constant provides a conversion between the mass of one mole of an entity and the average atomic mass of that same entity. N_A atoms with an average atomic mass of x u will have a mass of x g.
5. A chemist can measure the mass of samples using a balance. Because the chemist knows the molar mass of substances, they can use this knowledge to calculate the number of particles in that sample of substance.
6. (a) Numerically, molar mass and average atomic mass are the same. The units are different: g for molar mass and u for average atomic mass. The molar mass is the average atomic mass in grams multiplied by the Avogadro constant. The mole is defined as an amount of substance that has the same number of particles as exactly 12 g of carbon-12.
   (b) convenience
7. (a) for a metallic element, the molar mass is the mass of N_A atoms of that element.
   (b) for a diatomic element, the molar mass is equal to the mass of 2 x N_A atoms of the element.
   (c) for a compound, the molar mass is equal to the combined mass of N_A atoms for each atom found in the compound.

Answers to Inquiry Questions

9.

Isotope	Percentage Abundance	Approximate Atomic Mass (u)
Ar-36	0.34	36
Ar-38	0.06	38
Ar-40	99.6	40

aam = [0.34(36.0) + 0.06(38.0) + 99.6(40)] u / 100 =[12.24 + 2.28 + 3984] u / 100 = 39.98u or 40 u

10.

Isotope	Percentage Abundance	Atomic Mass (u)
Ga-69	60.0	69
Ga-71	40.0	71

aam = [60(69.0) + 40(71.0)] u / 100 = [4140 + 2840] u / 100 = 70 u

11.

Isotope	PercentageAbundance	Atomic Mass (u)
Ge-70	20.5	70
Ge-72	27.4	72
Ge-73	7.8	73
Ge-74	36.5	74
Ge-76	7.8	76

aam = [(20.5)(70) + (27.4)(72) + (7.8)(73) + (36.5)(74) + (7.8)(76)] u / 100
aam = [1435 + 1972.8 + 569.4 + 2701 + 592.8] u / 100 = 72.7u

12.

Isotope	Percentage Abundance	Atomic Mass (u)
K-39	x	39.0
K-41	100 – x	41.0

Given that the *aam* = 39.1 u
39.1 u = [*x*(39.0) + (100 *x*)(41.0)] u / 100
3910 = [39.0*x* + 4100 – 41.0*x*]
2.0*x* = 190 so *x* = 95.
Thus natural potassium is 95% K-39 and 5% K-41.

13. In each case, divide by the molar mass. ($n = m / M$).
(a) $n = 0.453 / 159.7 = (2.84)(10^{-3})$ mol.
(b) $n = 50.7 / 98.09 = 0.517$ mol.
(c) $n = (1.24)(10^{-2}) / 152.00 = (8.15)(10^{-5})$ mol.
(d) $n = (8.2)(10^{2}) / 187.37 = 4.38$ mol.
(e) $n = 12.3 / 97.95 = 0.126$ mol.

14. In each case, divide by the Avogadro constant ($n = N / N_A$).
(a) $n = (4.27)(10^{21}) / (6.02)(10^{23}) = (7.09)(10^{-3})$ mol.
(b) $n = (7.39)(10^{23}) / (6.02)(10^{23}) = 1.23$ mol.
(c) $n = (5.38)(10^{22}) / (6.02)(10^{23}) = (8.94)(10^{-2})$ mol.
(d) $n = (2.91)(10^{23}) / (6.02)(10^{23}) = 0.483$ mol.
(e) $n = (1.62)(10^{24}) / (6.02)(10^{23}) = 2.69$ mol.
(f) $n = (5.58)(10^{20}) / (6.02)(10^{23}) = (9.27)(10^{-4})$ mol.

15.

Isotope	Molar Mass (g/mol)	Sample Mass (g)	Number of Molecules	Number of Moles of Molecules	Number of Moles of Atoms
NaCl	58.44	58.44	$(6.02)(10^{23})$	1.00	2.00
NH_3	17.04	24.8	$(8.79)(10^{23})$	1.46	5.84
H_2O	18.02	1.58	$(5.28)(10^{22})$	$(8.77)(10^{-2})$	$(2.63)(10^{-1})$
Mn_2O_3	157.88	10.5	$(4.00)(10^{22})$	$(6.64)(10^{-2})$	0.332
K_2CrO_4	194.20	$(9.67)(10^{-1})$	$(3.00)(10^{21})$	$(4.98)(10^{-3})$	$(3.49)(10_{-2})$
$C_8H_8O_3$	152.16	$(1.99)(10^{3})$	$(7.90)(10^{24})$	13.1	249
$Al(OH)_3$	78.01	$(6.66)(10^{4})$	$(5.14)(10^{26})$	$(8.54)(10^{2})$	$(5.98)(10^{3})$

16. (a) $PtBr_2$: M = 195.08 +2 (79.90) = 354.88 u.
(b) $C_3H_5O_2H$: M = 3(12.01) + 5(1.01) + 2(16.00) + 1.01 = 74.09 u.
(c) Na_2SO_4: M = 2(22.99) + 32.07 + 4(16.00) = 142.05 u.
(d) $(NH_4)_2Cr_2O_7$: M = 2[14.01 + 4(1.01)] + 2(52.00) + 7(16.00) = 252.10 u.
(e) $Ca_3(PO_4)_2$: M = 3(40.08) + 2[30.97 + 4(16.00)] = 310.18 u.
(f) Cl_2O_7: M = 2(35.45) + 7(16.00) = 182.90 u.

17. (a) Sample = 3.70 mol = *x* g
Mole = 1.0 mol = 18.02 g
3.70 / 1.0 = *x* / 18.02
x = 66.7
There are 66.7 g of water.

1 mole H_2O
$(6.02)(10^{23})$ molecules
18.02 g

(b) Sample = $(8.43)(10^{23})$ molecules = x g
Mole = $(6.02)(10^{23})$ molecules = 239.20 g
$(8.43)(10^{23})$ / $(6.02)(10^{23})$ = x / 239.20
x = 335
There are 335 g in the sample.

1 mole PbO_2 $(6.02)(10^{23})$ molecules 239.20 g

(c) Sample = 14.8 mol = x g
Mole = 1.0 mol = 253.33 g
14.8 / 1.0 = x / 253.33
x = $(3.75)(10^3)$
There are $(3.75)(10^3)$ g in the sample.

1 mole $BaCrO_4$ $(6.02)(10^{23})$ molecules 253.33 g

(d) Sample = $(1.23)(10^{22})$ molecules = x g
Mole = $(6.02)(10^{23})$ molecules = 70.90 g
$(1.23)(10^{22})$ / $(6.02)(10^{23})$ = x / 70.90
x = 1.45
There are 1.45 g in the sample.

1 mole Cl_2 $(6.02)(10^{23})$ molecules 70.90 g

(e) Sample = $(9.48)(10^{23})$ molecules = x g
Mole = $(6.02)(10^{23})$ molecules = 36.46 g
$(9.48)(10^{23})$ / $(6.02)(10^{23})$ = x / 36.46
x = 57.4
There are 57.4 g in the sample.

1 mole HCl $(6.02)(10^{23})$ molecules 36.46 g

(f) Sample = $(7.74)(10^{19})$ molecules = x g
Mole = $(6.02)(10^{23})$ molecules = 159.70 g
$(7.74)(10^{19})$ / $(6.02)(10^{23})$ = x / 159.70
x = $(2.05)(10^{-2})$
There are $(2.05)(10^{-2})$ g in the sample.

1 mole Fe_2O_3 $(6.02)(10^{23})$ molecules 159.70 g

18. C_6H_6: M = 78.12 g/mol and given m = 45.6 g
$n = m / M$ = 45.6/78.12 = 0.584 mol
Using $n = N / N_{A,}$ the number of molecules is
$N = (0.584)(6.02)(10^{23}) = (3.52)(10^{23})$ molecules.
19. Since there are 12 atoms per molecule, there are $12(3.52)(10^{23}) = (4.22)(10^{24})$ atoms.
20. (a) for one atom, mass = 131.29 u.
(b) for one mole, mass = 131.29 g.
(c) 131.29 g = $(6.02)(10^{23})(131.29\text{u})$ or 1 g = $(6.02)(10^{23})$ u
so 1 u = $[1.00 / (6.02)(10^{23})]$ g = $(1.66)(10^{-24})$ g.
(d) Mass of one atom is 131.29 u = $131.29(1.66)(10^{-24})$ g = $(2.18)(10^{-22})$ g.
(e) Mass of one mole is 131.29 g = $131.29(6.02)(10^{23})$ u = $(7.90)(10^{25})$ u.
21. First find the number of moles of C atoms by addition: n = 1(0.237) + 2(2.38) = 5.00 mol; then the number of atoms is $5.00(6.02)(10^{23})$ or $(3.01)(10^{24})$.
22. Consider the water first. In $(3.49)(10^{23})$ molecules there are $2(3.29)(10^{23})$ or $(6.98)(10^{23})$ H atoms.
Now consider CH_3OH. Using $n = m / M$, the number of moles of molecules is 78.1/32.05, or 2.44 moles.
Since each molecule contains four atoms of H, there are 4(2.44) or 9.76 moles of H atoms.
This is $9.76(6.02)(10^{23})$ or $(5.88)(10^{24})$ H atoms.
The total number of H atoms is $(6.98)(10^{23}) + (5.88)(10^{24}) = (6.58)(10^{24})$ atoms.

23. There are two nitrate ions per formula unit of the compound.
Thus there are $2(3.76)(10^{-1}) = (7.52)(10^{-1})$ moles of nitrate ions.
Thus there are $(7.52)(10^{-1})(6.02)(10^{23}) = (4.53)(10^{23})$ nitrate ions
24. The required mole ratio is oxygen/ethanol = 3 / 1. Find the number of moles of ethanol. $n = m / M = 92.0 / 46.08 = 2$ mol. Thus 3(2) = 6 mol of oxygen molecules are required. Thus 6(32) g or 192 g of oxygen are required.

25.

Isotope	Percentage Abundance	Atomic Mass (u)
Br-79	x	79.0
Br-81	100 – x	81.0

From the Periodic Table, *aam* = 79.9 u.
$79.9 \text{ u} = [x(79.0) + (100 - x)(81.0)] \text{ u} / 100$
$7990 = [79.0x + 8100 - 81.0x]$
$2.0x = 110$, so $x = 55.0$
Natural bromine is 55.0% Br-79 and 45.0% Br-81.

26. (a) $NaCl + AgNO_3$
One formula unit reacts with one formula unit, so one mole reacts with one mole.
(b) To solve, find the molar masses and use ratio and proportion.

NaCl	$AgNO_3$
58.44 g	169.88g
29.2 g	*x* g

Thus 58.44 / 29.2 = 169.88 / *x* and *x* = 84.9. Therefore 84.9 g of silver nitrate will react with 29.2 g of sodium chloride.

27. The relative masses of atoms must be the same on all planets so we can use the values on our Periodic Table.

Entity	Number of Wogs	Mass in Wibbles
C-12	1.0	12.0
(a) N	1.0	14.01
(b) O_2	$(5.00)(10^{-1})$	$(5.00)(10^{-1})(32.00) = 16.0$

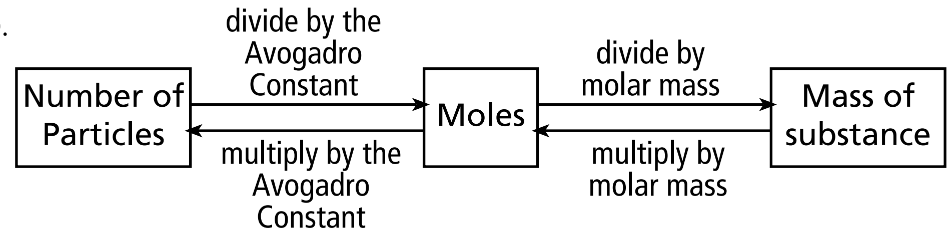
(c) 1 wog = $(2.50)(10^{21})$ entities.
Converting to moles, this is $(2.50)(10^{21}) / (6.02)(10^{23}) = (4.15)(10^{-3})$,
so 1 wog = $(4.15)(10^{-3})$ mol.
1 mol of H atoms has a mass of 1.01 g,
then $(4.15)(10^{-3})$ moles have a mass of $(4.15)(10^{-3})(1.01)\text{g} = (4.19)(10^{-3})$ g,
thus 1 wog of H atoms has a mass of $(4.19)(10^{-3})$ g.

Answers to Communication Questions

28. The Avogadro constant is defined as the number of entities in a mole. A mole is defined as an amount of matter having as many entities as there are atoms in 12 g of carbon-12. This definition does not tell us what the number of entities is. We must find an experiment that allows for the indirect counting of the number of entities in a mole.

29. The atomic mass scale was defined using C-12 having mass 12 u. Since proton and neutron masses are not exactly 1 u, — and electrons have mass — all other masses are not whole numbers.

30.



31. Avogadro's constant is applied to atoms, elements, molecules, compounds, formula units, or ions. N_A describes the relationship between a single particle and 6.02×10^{23} particles. This conversion only applies when moving between the same particle, not different particles. Carbon dioxide is a molecule that contains three atoms. One mole of carbon dioxide contains 6.02×10^{23} molecules of carbon dioxide, or $(3)(6.02 \times 10^{23}) = 1.8 \times 10^{24}$ atoms.

Answers to Making Connections Questions

32. (a) Molar mass of Fe = 55.85 g/mol
    number moles of Fe required per day = 0.0148 g ÷ 55.85 g/mol = 2.65×10^{-4} mol Fe
    (b) There is one mole of Fe atoms per molecule of ferrous gluconate, therefore the number of moles of iron taken per day is equal to the number of moles of ferrous gluconate required per day.
    Molar mass of $Fe(C_6H_{11}O_7)_2$ = 446.33 g/mol
    Number of moles of $Fe(C_6H_{11}O_7)_2$ required per day = 2.65×10^{-4} mol
    Mass of $Fe(C_6H_{11}O_7)_2$ required = (446.33 g/mol)(2.65×10^{-4} mol) = 0.1182 g, or 118.2 mg
    Therefore 118.2 mg of ferrous gluconate is required per day.
    (c) For information regarding the addition of elemental iron to cereals, check out Health Canada's web site and Food Guide for more information: http://www.hc-sc.gc.ca.
33. Molar mass of niacin = 123.14 g/mol
    Molar mass of vitamin B (niacinamide) = 122.15 g/mol.
    (a) Number of moles of niacinamide in a vitamin supplement = 0.100 g ÷ 122.15 g/mol = 8.19×10^{-4} mol
    Equivalent mass of niacin = (8.19×10^{-4} mol)(123.14 g/mol) = 0.10085 g = 100.9 mg
    (b) The recommended nutrient intake (RNI) for niacin is:
    average adult male: 14 – 20 mg
    average adult female: 10 – 14 mg
    (c) Foods high in niacin include tuna, wheat bran, and Ovaltine.
    (d) Symptoms of a niacin overdose include: blurred vision, excessive itching, nausea, abdominal pain, and lightheadedness. Overdose is usually not life threatening but in any case, the nearest poison control center should be called.
    (e) More details on dietary inputs and vitamins can be found at the following sites: http://webmd.lycos.com and http://healthyeating.org/books/foodfacts/html/data/data-fs.html.