Unit 5 – Counting Particles Too Small to See

Instructional goals

1. Use Avogadro’s Hypothesis and experimental data to determine the relative mass of molecules.

2. Use experimental data to determine the number of items in a sample without actually counting them.

3. Use experimental data to determine the molar mass of elements.

4. Given the chemical formula of a substance, determine the molar mass.

5. Given the mass of a substance, determine
   a. the number of moles of the sample
   b. the number of atoms or molecules in the sample

6. Given the number of moles of a substance, find
   a. the mass of the sample
   b. the number of atoms or molecules in the sample

7. Given the formula of a compound, determine its % composition.

8. Given data about the % composition of a sample, determine the empirical formula of the compound.

9. Given the empirical formula and information about the molar mass of the compound, determine the molecular formula.

Sequence

1. Discussion: counting by weighing, Avogadro’s Hypothesis

2. Activity: Relative mass of hardware, data collection and evaluation

3. Post-lab discussion – definition of the mole, molar mass as relative mass of elements, worksheet 1

4. Size of a mole worksheet

5. Use of molar mass as conversion factor between “how much” and “how many”, begin Empirical Formula lab, worksheet 2 – p1

6. Finish Empirical Formula lab, evaluation of data, comparison of class results

7. Post-lab discussion, worksheet 2 - p2

8. Empirical and molecular formulas – quiz on mass ↔ mole conversions
9. Percent composition, worksheet 3

10. Begin nail lab – unit review

11. Part 2 of nail lab - finish unit review

12. Unit test

**Overview**

One of the greatest challenges that early chemists faced was trying to find a way to connect the mass of a substance to the number of particles in the sample. The findings that “elementary particles” combined in fixed ratios by weight led Dalton to his “atomic model” of matter. However, knowing that the mass ratio of oxygen to hydrogen in water is 8:1 doesn’t tell you *how many* atoms of each element are involved unless one knows the relative mass of each kind of atom.

The trick in assigning relative masses to elements is to be able to feel confident that the samples you are comparing have the same number of particles. It would not do, for example, to try to compare the mass of a baseball to a golf ball by weighing backpacks filled with each kind of ball. As we have seen earlier in the experiment comparing mass and volume of different metals, we could not say with certainty whether the iron was more dense because it had more particles per given volume than a sample of aluminum, or that its individual particles were more massive than those of aluminum, or some combination of both.

In estimating atomic weights, Dalton was confronted with certain grave difficulties. Since it is impossible to weigh single atoms, any system of atomic weights must be formulated on a comparative basis. The atom of some element must be arbitrarily selected as the reference weight. Dalton chose the hydrogen atom and assigned *one* as its weight. The atomic weight of oxygen could then be found by either (1) comparing the weights of equal numbers of oxygen and hydrogen atoms or (2) finding by analysis the combining weights of oxygen and hydrogen in water. Dalton considered the first approach but rejected it. Since to him, atoms in a gas were analogous to a pile of shot, and since he believed that atoms of different gases varied in diameter, therefore equal volumes of gases could not contain equal numbers of atoms. The second approach he considered valid. However, reflection will show that it is valid only when the ratio of atomic combination is known.¹

The breakthrough came when Avogadro devised an hypothesis to account for the observations of Gay-Lussac regarding the reacting volumes of gases. Gay-Lussac had noted that gases appeared to react in simple integer ratios. It was observed that two volumes of carbon monoxide combined with one volume of oxygen to form two volumes of carbon dioxide, and ammonia contained nitrogen and hydrogen in a 1:3 ratio, and so on. Gay-Lussac’s law of combining volumes suggested that equal volumes of gases, at the same temperature and pressure, contained equal numbers of particles. However, there were problems. Two volumes of hydrogen gas combined with one volume of oxygen gas to produce *two* volumes of steam. If H combines with O in a 2:1 ratio and each volume contained *the same number of* particles of gas, one would expect only *one* volume of steam to be formed.

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Avogadro’s paper of 1811, based on Gay-Lussac’s law and Dalton’s atomic theory, reconciled the … problems. Starting with the assumption that equal volumes of all gases contain equal numbers of molecules under similar conditions, Avogadro proceeded to analyze the facts of gaseous combination. Using the examples discussed by Gay-Lussac, he showed that the ambiguities disappeared if he assumed that the molecules involved in typical reactions might split into “half-molecules”; that is, he supposed the existence of molecules of elemental gases which contained more than a single atom. He did not use the term “atom”, but always used the term “half-molecule” as its equivalent.(Ihde)

Once we accept Avogadro’s Hypothesis, we can compare the mass of various gases and deduce the relative mass of the molecules. From there we can pick a weighable amount of the lightest element (say 1.0 g), then use mass ratios to assign the masses of the other elements. How many atoms are present in the sample? As unlikely as it might seem, it’s really not important to know. For if two volumes of hydrogen combine with one volume of oxygen gas, it’s reasonable to assume that two molecules of hydrogen are reacting with each molecule of oxygen. In fact, the word chosen to represent the standard weighable amount of stuff, the mole, comes from the Latin: lump of stuff.

For the purposes of Dalton’s theory their absolute sizes and weight were irrelevant. It was not necessary to know (e.g.) how many atoms there are in a gram of hydrogen, so long as one could assume that this number was very large indeed. For all one could infer by the application of Dalton’s ‘General Rules’ were the relative weights of atoms and molecules of different substances. One could establish that each oxygen atom weighed 4/3 times as much as each carbon atom – but there was no way of knowing whether they weighed 4 and 3 millionths, or 4 and 3 billionths, or 4 and 3 trillionths of a gram. The ‘atomic weights’ and ‘molecular weights’ of chemistry have always been expressed as multiples of an arbitrary unit – the weight of one hydrogen or carbon or oxygen atom being fixed by definition as ‘1’ or ‘12’ or ‘16’.²

The unit builds on the use of Avogadro’s Hypothesis introduced in unit 4. However, instead of using the hypothesis to deduce formulas of compounds, we now use it to deduce the relative masses of elements. Unlike in solids, the particles of a gas are far enough apart that their size has little effect on the volume. This discussion leads to the Relative Mass activity which provides students a tangible example of comparing masses of the same number of particles in the gaseous state. The lab and ensuing discussion sets the stage for the definition of the mole as a weighable amount of an element or compound. In worksheet 1 students use the densities of hydrogen and oxygen to determine the relative mass of oxygen. Then they examine how mass composition could be used to determine the molar masses of elements. Students are introduced to some examples to help them to get some sense of the magnitude of Avogadro’s number.

Molar mass is introduced as a factor relating how much (mass) to how many (moles). Students practice performing these calculations both in the Empirical Formula lab and by working through worksheet 2. Students also learn how to determine the molecular formula of a compound from the empirical formula and molar mass. Finally, they learn how to determine the percent composition of elements in compounds.

Avogadro’s number is introduced as an empirically determined number after students have understood the concept of molar mass as the relative mass of one ‘standard lump’ of particles. The size of this number serves to remind us both how tiny these particles are and that their size is knowable – reinforcing what the students observed about particle size in Unit 1. Students learn to use Avogadro’s number as a conversion factor between moles and numbers of particles, but the emphasis of the unit is on the mass-count relationship of molar mass and how we can use it.

Instructional Notes

1. Discussion: counting by weighing, Avogadro’s Hypothesis

Bring out a large bag of Styrofoam packing peanuts and ask students if they would like the task of counting how many peanuts were in the bag. Then, inform them that you had found that 100 peanuts had a mass of 5.5 g. Ask again if they could think of a way to determine the number of peanuts presenting the sample. Someone should be able to suggest that we could measure the mass of the bag of peanuts and divide by the mass of the 100 peanuts to get a reasonable estimate of the number on the bag.

Explain that this was the challenge that early chemists faced when they wanted to get some sense of how many atoms or molecules were present in a sample they were investigating. Remind students that back in Unit 1 we found that iron was more dense than aluminum. Two possible models arose to account for this difference.

A. The masses of Al and Fe atoms are about the same, but there are more atoms of Fe than atoms of Al in each cm$^3$ sample.

B. One cm$^3$ samples of Fe and Al contain about the same number of atoms, but the Fe atoms are more massive.

A third possibility – that both the size and the mass of the atoms of these two elements were different – also came up. At the time, we did not have enough evidence to make a decision about these possible models.

While the reason for density variation between particle types is difficult to determine for liquids and solids, a conclusion can be reached more easily for gases due to the fact that particles in a gas are widely spaced. This means that particle size does not have an effect on the volume that a given number of gaseous particles occupy. Review the experimental evidence of Gay-Lussac’s combining volumes and Avogadro’s Hypothesis. This hypothesis not only allows us to deduce the numbers of molecules that react with one another based on volumes, but also enables us to determine the relative masses of elements and compounds.

2. Activity: Relative mass

Apparatus
Snap-cap vials or film canisters
Equal numbers of pieces of some similar kind of item (beads, seeds, metal shot, hardware, etc)
Balances

Pre-activity discussion
Bring out a set of vials containing the items you have prepared. Ask the students to pretend that the items in the vials are enlarged particles of various gases. Since the vials are the same size, we will assume that each contains the same number of particles. Their job is to determine the relative mass of each kind of item in the vial, without opening the vial to examine single items.
3. Post-activity discussion – definition of the mole, molar mass - worksheet 1

In the first exercise in worksheet 1, students should conclude that each molecule of oxygen is 16x as massive as a molecule of hydrogen. Once early chemists realized they could determine the relative masses of molecules of gases, they chose weighable samples of the elements as the standard amount of each element. While 12.00 g of C\textsuperscript{12} was eventually chosen as the standard weighable amount, a mole (from the Latin - “lump of stuff”), it was originally defined as 1.0g of the lightest element known – hydrogen.

Weigh out 1.0g of hydrogen atoms and, by definition, you have a mole of atoms. How many atoms are in that sample? The answer to that question would not be determined for many years. But one didn’t need to know that value. If one compared the masses of equal volumes of hydrogen and another element, at the same temperature and pressure, then the relative mass of that element would be the mass of one mole of that element. This is true whether the sample contained a million atoms, a billions atoms or 6.02 x 10\textsuperscript{23} atoms. The molar masses (formerly known as atomic weights) we find in the Periodic Table of the elements were originally determined relative to the mass of hydrogen.

In the second part of the worksheet students see that they didn’t always have to compare masses of gaseous elements to hydrogen at the same temperature and pressure. The percent composition of oxides could be used to determine molar masses of some elements. It is important to point out that the assumption that the atoms combine in a 1:1 ratio may not always be true. If (as Dalton believed until his death) the formula for water was HO, then the mass ratio in water implies that oxygen atoms are 8x as massive as hydrogen. Yet, as we saw in step 1, oxygen gas is 16x as dense as hydrogen gas. Since volume ratios lead us to believe that water has 2 atoms of hydrogen for each atom of oxygen, one must make an adjustment to have a fair comparison. To help students see this, show students a diagram like the one at right. Ask students if they think it is fair to conclude from the scale readings that each atom of oxygen is \( \frac{100}{12.5} = 8.0x \) as heavy as an atom of hydrogen. They should respond that one should either cut the mass of hydrogen in half or double the mass of oxygen. Either approach leads one to the correct conclusion that an atom of oxygen is 16x as massive as a hydrogen atom. Since we chose 1.0 g to be the mass of one mole of hydrogen, then the molar mass of oxygen must be 16.0 g.

In step 3 students simply double Dalton’s values from column 3 (with the exception of H). This approach yields values that agree with the molar masses in the periodic table until one gets to silver oxide. There, the value calculated for silver (216 g) is twice the accepted value. Ask students to account for this discrepancy. Ideally, they should conclude that there are two atoms of silver for every atom of oxygen in this compound.

4. Size of a mole worksheet

A number of examples are given to show that a mole contains a number of particles that is difficult to imagine. After they have completed the calculations, ask the students how large a mole of water would be. Show a graduated cylinder containing 18 mL of water. This should help them appreciate just how small molecules really are.
5. Molar mass as conversion factor, begin empirical formula lab, worksheet 2 – p1

**Apparatus**

**For each lab group:**
- 3-8 pieces of zinc metal\(^3\) or samples of mossy zinc
- 50 mL 3M HCl
- 100 or 150 mL beaker
- burner and stand

**for the class:**
- fume hood to exhaust HCl fumes
- several hot plates

**Pre-lab discussion**

Demonstrate the reaction between a piece of zinc metal and hydrochloric acid. Inform them that one product of the reaction is a compound of zinc and chlorine, which remains dissolved in the solution. You can collect a sample of the gas evolved and demonstrate that it behaves like hydrogen. Their task is to determine the empirical formula – the simplest ratio of atoms in the compound obtained by experimental evidence. Ask the students how they could obtain just the solid product zinc chloride. Some should suggest that one could evaporate the water once the reaction was complete. Distribute copies of the lab handout and assign different numbers of pieces of zinc to different groups. If you use mossy zinc, you should have pre-weighed samples ranging from 1.5 – 4.0 g available to the students so they will find that the mass ratio does not depend on the starting mass of Zn. Make sure that the mass of zinc in the largest sample does not exceed \(0.050 \, \text{L} \times \frac{3.0 \, \text{moles HCl}}{1 \, \text{L}} \times \frac{1 \, \text{mole Zn}}{2 \, \text{moles HCl}} \times \frac{65.4 \, \text{g}}{1 \, \text{mole}} = 4.91 \, \text{g}\) of the metal.

**Performance notes**

Once the students have carefully massed the Zn, place the pieces in the beaker and added the 3M HCl, they should observe the reaction for a while, then place the beaker on the hot plate in the fume hood or designated place. While the reaction is proceeding, they should begin worksheet 1. This worksheet gives students opportunities to practice molar mass and the number of particles per mole as factors that allow one to relate one kind of measure (how much, how many lumps, how many individual items) to one another.

The reaction should be complete by the end of the class period. However, you will need to babysit the beakers as you evaporate most of the remaining HCl solution in the beaker.

6. Finish empirical formula lab

The following day, allow the students to obtain their labeled beakers and observe the contents. The zinc chloride is hygroscopic, so even if you heated the sample to dryness the day before, it will be somewhat soupy now. Have the students heat the sample to drive off the remaining water, allow the beaker to cool to the touch before placing it on the balance to find the mass of the beaker and zinc chloride. To be sure that all the water has been driven off, they should heat the sample more strongly this time; be sure to not heat so strongly that the sample begins to smoke. Once cool, they should mass the beaker and product again. If the two masses agree to within 0.02 g, they are finished. If not, they should heat the beaker for a third time.

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\(^3\) These can be purchased as \(\frac{1}{2}\)” pieces or cut from a sheet of zinc.
Elicit from the students how they could go about finding the mass of chlorine present in their sample of zinc chloride. Hopefully, someone will suggest that they could subtract the mass of zinc from the mass of the product to obtain the mass of the chlorine. They now have enough information to determine the moles of zinc and chlorine present in their compound and to calculate the mole ratio of chlorine to zinc. This will allow students to determine the empirical formula of the compound. If time allows, have the students post their values for moles of zinc, moles of chlorine and the mole ratio.

7. Post-lab discussion, worksheet 2 - p2
   Once the groups have posted their values, see if they can suggest the empirical formula for zinc chloride. Does the empirical formula agree with the formula they would expect from the charges of the ions? Ask students to explain how they could have all obtained nearly the same value despite the fact that they used different amounts of zinc.

8. Empirical and molecular formulas – quiz on mass ↔ mole conversions
   Once students know how to calculate the empirical formula of a compound, show examples of how they can check to see if this formula represents the actual number of atoms of each type in the compound. To find out, they must compare the molar mass, obtained by some other means, (mass compared to known gas, or from colligative properties) to the formula mass.

9. Percent composition, worksheet 3
   This worksheet gives students practice at finding % composition and using this information to determine the empirical and molecular formulas of a compound.

10. Unit review

11. Unit test