Figure 1.6 shows a high-energy particle accelerator. Such an instrument helps modern scientists probe inside tiny particles of matter by splitting them apart. Using this tool, scientists can even create new particles that do not exist in nature. Work of this kind is based on modern theories that explain the composition of matter. These theories, in turn, are rooted in the work and the ideas of scientists from centuries ago. These scientists used techniques and tools that are “low-tech” by today’s standards. Even so, early chemists devised atomic theories that are still useful today.

**Early Observations**

During the 1600s and 1700s, scientists improved laboratory techniques for isolating pure substances and analyzing their properties. The scientists gathered a great deal of information about specific substances and the ways that they interact. In some cases, scientists observed an action or condition so consistently that they were convinced it would always happen. When scientists are convinced of the regularity of certain observations, they generalize their observations as scientific laws. Several of these early laws could be explained by the hypothesis that matter is made up of tiny particles. In this section, you will discover how that hypothesis developed to become modern atomic theory.

**Dalton’s Atom**

John Dalton (1766–1844) was an English scholar and teacher. He published a comprehensive atomic theory in 1808. The heart of Dalton’s theory was that every substance is made up of indivisible atoms. Further, the key difference between atoms of different elements is their mass. Dalton’s theory is summarized below. Note how well it explained many observations and laws.

For example, by the late 1700s, scientists knew that when substances react, the total mass of the substances before and after the reaction is always the same. If matter is made up of indestructible particles, this law makes sense. Particles are only rearranged during reactions. They are not destroyed or created. Today, chemists still use many parts of this theory to explain the behaviour of matter. Figure 1.7 shows how Dalton might have pictured the atom.

**Dalton’s Atomic Theory**

- All matter is made up of small particles called atoms.
- Atoms cannot be created, destroyed, or divided into smaller particles.
- All atoms of the same element are identical in mass and size, but they are different in mass and size from the atoms of other elements.
Compounds are formed when atoms of different elements combine in fixed (definite) proportions. The tiniest particles of any compound always contain the same types and relative numbers of atoms.

Chemical reactions change the way atoms are grouped, but the atoms themselves are not changed in reactions.

Many of Dalton’s conclusions were based on assumptions. If the assumption was wrong, it led to an incorrect conclusion. For example, Dalton assumed that atoms combined in the simplest possible way. He knew that water contained hydrogen and oxygen, so he proposed that the formula for water was OH. He assigned hydrogen a mass of one unit. Then, Dalton used measurements made by French chemist Joseph Proust (1754–1826). According to Proust’s measurements, water contains eight times more oxygen by mass than hydrogen. Based on Dalton’s assumption, an atom of oxygen would then have a mass of eight units. Dalton developed a system of symbols, shown in Figure 1.8, to keep track of his assumptions about how atoms combined.

Dalton’s assumptions about the composition of water, and several other compounds, were inaccurate. As a result, some of his calculations of relative masses were also inaccurate. However, Dalton used his theory to predict different ways in which a given pair of elements might combine. He proposed, for example, that there should be several different compounds containing nitrogen and oxygen. These included NO, N2O, and NO2. When his prediction was verified experimentally, doubts about Dalton’s atomic theory gave way to widespread acceptance.

Electricity and the Atom

Sometimes science fosters technological discoveries, and sometimes new technology stimulates scientific discoveries. A technological achievement that helped scientists improve on Dalton’s theory was the refinement of the gas discharge tube. A gas discharge tube is a sealed glass vessel that contains a gas at low pressure. As electricity flows through the gas, a “ray” is formed across the length of the tube, and light is produced. The rays produced in gas discharge tubes are called cathode rays. In 1855, Heinrich Geissler (1814–1879), a German glass-blower and mechanic, improved the gas discharge tube. Figure 1.9 shows a modern version of a Geissler gas discharge tube.

Did You Know?

John Dalton was colour-blind to red, a condition that made it difficult for him to describe and identify chemicals by sight. In 1794, Dalton wrote about colour-blindness. His paper was the earliest scientific description of this condition. Colour-blindness is still sometimes referred to as Daltonism.

TRETCH Your Mind

Some ancient Greek philosophers speculated that the universe must be composed of small particles that could not be broken down. They used the Greek word atomos ("indivisible") to describe these particles, which were thought to be separated by empty space. Ancient Greek thinkers arrived at their ideas by a series of logical arguments. They did not use experimental investigation to test or develop their ideas. Would you consider the methods of the ancient Greek philosophers to be scientific? Explain why or why not.
Evidence of Electrons
Several experiments with gas discharge tubes led researchers to infer that matter contains tiny particles that have negative charges. This inference may seem obvious today. At the end of the nineteenth century, however, many scientists were reluctant to abandon the central theme of Dalton’s useful atomic theory. They did not want to believe that Dalton’s indivisible atoms might actually be made up of even smaller particles.

Develop a Theory
As you are discovering, scientists developed theories about the structure of the atom without ever seeing the atom. They used models to represent their theories visually. In this activity, you will construct a mystery box and develop your own model to show what is inside it. Then you will challenge a partner to collect evidence and develop a theory about what is inside.

Materials
- cardboard box, the size of a shoe box
- objects to place inside box
- adhesive tape
- thin, stiff wire

Procedure
1. Design a mystery box. Keep in mind that:
   - a simple but creative mystery box is better than one that is too complicated
   - your box may not contain any liquid that could spill or any substance that could decompose, such as food
   - your design must allow for simple tests or experiments, such as probing with a thin wire or shaking
2. Construct your box. You can
   - put in one or two objects that can move and make noise when the box is tilted
   - tape a few objects to the inside of your box
3. Draw a model of the inside of your box. Your model must be based on the inferences you think your partner can make about it. Do not show your model to your partner.

Find Out
4. Seal your box, and exchange boxes with your partner.
5. Perform simple tests to determine what is inside your partner’s box. You may not open the box. Make a table like the one below to record the tests you performed and what you can infer about the internal structure of the box. Give your table a title.

<table>
<thead>
<tr>
<th>Tests Conducted on Box</th>
<th>Observations and Evidence Collected</th>
<th>Inferences Made Based on Evidence</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

6. Put your inferences together to develop a theory of the internal structure of the box. Then draw a model of the inside of the box.

What Did You Find Out?
1. Compare your model with your partner’s. How similar are they? Which inferences could account for the differences between them?
2. Which test yielded the most useful evidence?
3. Having seen your partner’s model of his or her box, what test did you not carry out that might have yielded useful results?
4. Suppose you are granted access to an X-ray machine to conduct further tests on the box. What are your hypotheses for the X-ray machine experiment? What predictions will you make before you begin?
In 1894, English physicist J. J. Thomson (1856–1940) used a new version of the gas discharge tube to obtain direct evidence that cathode rays were actually a stream of negatively charged particles. Thomson’s modified gas discharge tube used charged plates to bend cathode rays around a curved path. He knew that by measuring the radius of their path, he could calculate information about the mass and charge of the particles. Figure 1.10 shows a simplified view of Thomson’s experiment.

Thomson was indeed able to work out a quantitative relationship between the charge and the mass of the negatively charged particles. He showed that either they had far more charge than any other particle then known, or they were much less massive than an atom. Later experiments confirmed the second option. These stable, negatively charged particles are now known as electrons. Each electron has less than \( \frac{1}{2000} \) the mass of a single hydrogen atom.

**Did You Know?**

J. J. Thomson once said, “At first there were very few who believed in the existence of these bodies smaller than atoms. I was even told long afterwards by a distinguished physicist who had been present at my lecture at the Royal Institution that he thought I had been ‘pulling their legs.’”

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**Figure 1.10** Based on his experiments with cathode ray tubes, Thomson concluded that matter contained tiny, charged particles. How would diagram (C) change if the positions of the positive plate and the negative plate were reversed?
Did You Know?

Thomson is sometimes called "the father of the electron." When he published his results in 1897, however, he referred to the cathode ray particles as corpuscles. It was not Thomson, but another scientist, G. Johnstone Stoney, who invented the name "electron" in 1891 to describe a unit of charge in electrolysis experiments. A third scientist, George Fitzgerald, argued that this electron and Thomson’s corpuscle were really the same thing.

Electrons and the Atom

Electrons were much less massive than atoms. Electrons also appeared to be present in all samples of matter. These discoveries suggested to scientists like J. J. Thomson that every atom contained electrons. Electrons are negatively charged, however, and samples of matter normally have no overall charge. Therefore, each atom would also have to contain a source of positive charge. Yet, it was not clear where the positive charge was. Thomson at first supported a model first suggested in 1902 by another English physicist with the same last name. This person was William Thomson, better known as Lord Kelvin (1824–1907). According to Kelvin’s model, atoms consisted of electrons embedded in a spherical cloud of positive charge.

Figure 1.11 shows a diagram representing the Kelvin/Thomson model. This model eventually became known simply as the Thomson model.

The electrons in this model are like raisins in a plum pudding or raisin’ bun. Thus, Thomson’s theory has been called the “plum-pudding” or “raisin-bun” theory. This model (also called the Thomson atom) could not account for a phenomenon that Thomson himself was studying. Radioactive elements had only recently been isolated in pure form. They appeared to be constantly emitting fast-moving, positively charged particles. These particles are called alpha particles and have about 7200 times the mass of an electron. The Thomson atom contained nothing similar to alpha particles, and gave no clues about how they might be formed.

Rutherford’s Experiment

In 1909, New Zealand-born physicist Ernest Rutherford (1871–1937) designed an elegant experiment to probe the structure of atoms. As shown in Figure 1.12 on the next page, Rutherford’s apparatus directed a stream of alpha particles from a shielded sample of radioactive polonium toward a very thin gold foil. Collisions with gold atoms, or parts of gold atoms, in the foil were expected to cause the alpha particles to change direction slightly and hit different parts of a fluorescent screen placed near the foil. Rutherford observed this deflection. He observed something else, as well. A small number of alpha particles bounced back from the gold foil. Rutherford did not expect this result.

Did You Know?

Ernest Rutherford won the 1908 Nobel Prize in Chemistry for finding that radioactive elements actually gave off three different types of emissions. These emissions are now called alpha particles, electrons, and gamma rays. He studied with J. J. Thomson at Cambridge, then taught at McGill University in Montreal from 1898 until 1907. Then he returned to England to develop his own research laboratory at the University of Manchester.
Based on his observations, Rutherford developed a new atomic theory. This theory included both electrons and positively-charged particles. The theory also explained the surprising alpha particle rebounds that he had observed. You can follow Rutherford’s logic by examining Figure 1.13.

**Did You Know?**

In 1904, a Japanese scientist called Hantaro Nagaoka proposed an atomic model that was similar to Rutherford’s. Nagaoka’s model described a disk-shaped atom with negatively charged atoms orbiting a positively charged nucleus. When Rutherford wrote about his atomic model in 1911, he noted that his results would be the same if Nagaoka’s model were correct.
Protons and a Nucleus

Rutherford’s atomic model is shown in Figure 1.14. Rutherford’s reasoning led him to propose that the atom had the following features:

- **A nucleus**: a central region that is positively charged, extremely small, and yet contains almost all of the atom’s mass. Such a nucleus would be far too dense to be a cloud or shell of positive charge, as Thomson had suggested. Rather, Rutherford visualized the nucleus as containing tiny, relatively massive particles, each with a single positive charge — protons. A proton is now known to have about the same mass as 1836 electrons.

- **Electrons**: particles with a single negative charge, located in the outer region of the atom. Electrons are much less massive than protons and neutrons. Rutherford’s model suggests that the electrons move around the nucleus rather like planets orbiting the Sun. Electrical attraction from oppositely charged protons in the nucleus would keep the electrons in orbit, just as the Sun’s gravity keeps planets in orbit.

- **Empty space**: a volume of space surrounding the nucleus that is very large, compared to the nucleus. The electrons exist in this space, but it is otherwise empty. Since the gold foil was a solid, Rutherford assumed that atoms in it were packed closely together. He reasoned that most alpha particles could pass through the foil due to space within individual atoms, not between them.

Evidence of the Neutron

Hydrogen is the element with the least-massive atoms. Rutherford hypothesized that hydrogen would have the simplest possible atom: one proton and one electron. It seemed logical that helium, the next-lightest element, would have two protons and two electrons. That would make a helium atom twice as massive as a hydrogen atom. According to experimental evidence, however, helium atoms are four times more massive than hydrogen atoms.

To explain these results, Rutherford hypothesized that there was a third subatomic particle in the atom. He hypothesized that this particle had the same mass as the proton, but no electrical charge. For this reason, he called the particle a neutron. A helium atom, then, would contain two protons and two neutrons (as well as two electrons), making it four times as massive as a hydrogen atom.

The neutron hypothesis could also explain the existence of isotopes. Isotopes are atoms of the same element that differ in mass but are chemically alike. All atoms of a given element were thought to contain the same number of protons, which accounted for their identical chemical properties. If atoms of the same element could have different numbers of neutrons, however, they would have different masses, as shown in Figure 1.15 on the next page.
Because they have no electric charge, neutrons were extremely difficult to isolate and study. In 1932 Rutherford and his colleague English physicist James Chadwick (1891–1974) gave clear experimental evidence of the existence of neutrons.

Evidence of Energy Levels

Two of the main objections to Rutherford's "solar-system" atom came from other physicists. According to physical theory, electrons moving around a nucleus should constantly emit energy in the form of light or radio waves. This process would cause the electrons to spiral into the nucleus, and the atom would collapse. Atoms, however, do not collapse. There is also no evidence that their electrons emit energy under normal conditions.

In a gas discharge tube, however, gases do emit light, but only when electrical energy is supplied to them. As well, each element emits specific colours of light. These colours are different for each element. The colours of light correspond to specific wavelengths. This means that the colours have different energies. Rutherford's atomic theory had no explanation for this behaviour. Figure 1.17 on the next page shows how instruments called spectroscopes separate light into different colours. Figure 1.18 shows the spectrum of light from hydrogen gas in a gas discharge tube.

The Danish physicist Niels Bohr (1885–1962) came up with an explanation for these observations. He hypothesized that electrons in an atom have certain allowed energies that enable the atom to remain stable. These allowed energies could be thought of as electron shells or energy levels. Electrons would be associated with specific energy levels. In addition, electrons could move only from one allowed energy level to another. They could not exist between the energy levels. By absorbing a specific quantity of energy, an electron could move to a higher energy level. By emitting the same quantity of energy, the electron could move back to its original energy level. Figure 1.19 shows how Bohr pictured electron energy levels. The diagram shows the energy levels in two dimensions for simplicity. In three dimensions, they would have the shape of spherical shells.
Instruments called spectroscopes separate light into different colours. Each coloured line in this hydrogen spectrum is produced by light with a certain energy.

Bohr proposed this new atomic theory in 1913. The theory fit very well with observations of light emitted from discharge tubes. It seemed reasonable that electrons in atoms of different elements would have different allowed energy levels, and would therefore absorb and emit light of different energies (different colours).

*The energy levels of electrons in an atom are the key feature of Bohr's theory.* Bohr showed how to derive mathematical equations that described these energy levels. The equations could be solved for the hydrogen atom (the simplest atom) by assuming that electrons moved in circular paths. As the electron absorbed energy, the equation showed that it could move farther from the positively charged nucleus. Bohr calculated the average distance of electrons in different energy levels from the nucleus of a hydrogen atom.

Atoms with more than one electron were too complex for Bohr to analyze mathematically in the same way as hydrogen. He did establish, however, that atoms with two or more electrons could have only a certain number of electrons in each energy level. This meant that atoms of each element would have a characteristic arrangement of electrons in different energy levels. Table 1.1 shows the maximum number of electrons that can occupy the first two energy levels.

### Table 1.1 Maximum Number of Electrons in First Two Energy Levels

<table>
<thead>
<tr>
<th>Energy level</th>
<th>Maximum number of electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>2</td>
</tr>
<tr>
<td>2</td>
<td>8</td>
</tr>
</tbody>
</table>

**Atomic Theory Evolves**

Current models of the atom are far more complex than Bohr's electron energy levels and Rutherford's nuclear model. For example, electron energy levels are now thought to be divided into sublevels. In many cases, electrons in the same energy level are grouped in pairs. In addition, scientists now believe that neutrons and protons are made of even smaller.
Presenting the Atom

It is often easier to understand a concept if you can see it represented visually. In this activity, your group will research an atomic theory. Your group will then use three methods of communication to explain its details to the rest of your class.

Materials

modelling materials, such as polystyrene balls, marbles, cardboard, modelling clay, and wire posterboard and drawing materials such as markers
computer three-dimensional modelling software (optional)

Procedure

1. Your teacher will divide the class into four groups. Each group will present an atomic theory according to one of the following scientists:
   - Dalton
   - Rutherford
   - Thomson
   - Bohr

2. Each group will learn about atomic theory according to Dalton, Thomson, Rutherford, or Bohr. As a group, you will communicate what you have learned in the following three ways:
   - a 10-minute presentation
   - a 3-D model or computer simulation showing how the scientist you researched might have pictured the atom

3. Use this textbook to do your research. As time permits, use other texts or Internet resources to supplement your research. The information you present should include:
   - information about the key experiment or experiments that provided evidence for the theory
   - details of the theory itself
   - explanation of why the theory needed to be modified (i.e., what observations did the theory fail to explain?)

What Did You Find Out?

1. Based on your group’s research and on the presentations of other groups, create a point-form summary of the development of atomic theory.

2. Each of the models of the atom that you have studied has some flaws. In other words, not all observations of matter can be explained using these models. Without knowing the details, do you think that the most modern atomic theory is likely to represent the best description of reality? Explain your answer.

3. Although Bohr’s atomic theory has flaws, it is still used today. Explain why a theory with known flaws is still taught and used.

particles called quarks. Instead of just three subatomic particles (neutrons, protons, and electrons), scientists have identified dozens of fundamental particles. It seems likely that with new observations and discoveries, theories of the atom will continue to evolve.
A Working Model of the Atom

Physicists and chemists have developed a complex and detailed model of the atom. A more simplified version explains many observations about chemicals and chemical changes. What are the key features of this model?

Protons and neutrons cluster together to form the central core, or nucleus, of an atom. For this reason, protons and neutrons are called nucleons. Electrons occupy the space that surrounds the nucleus of the atom. Table 1.2 and Figure 1.20 summarize the general features and properties of atoms and their three types of subatomic particles.

Table 1.2 Properties of Protons, Neutrons, and Electrons

<table>
<thead>
<tr>
<th>Subatomic particle</th>
<th>Relative charge</th>
<th>Symbol</th>
<th>Mass (in g)</th>
<th>Radius (in m)</th>
</tr>
</thead>
<tbody>
<tr>
<td>proton</td>
<td>1+</td>
<td>p⁺</td>
<td>1.67 × 10⁻²⁴</td>
<td>10⁻¹⁵</td>
</tr>
<tr>
<td>neutron</td>
<td>0</td>
<td>n⁻</td>
<td>1.67 × 10⁻²⁴</td>
<td>10⁻¹⁵</td>
</tr>
<tr>
<td>electron</td>
<td>1−</td>
<td>e⁻</td>
<td>9.02 × 10⁻²⁸</td>
<td>smaller than 10⁻¹⁸</td>
</tr>
</tbody>
</table>

An average atom is about 10⁻¹⁰ m in diameter. Such a tiny size is hard to visualize. If an average atom were the size of a grain of sand, a strand of your hair would be about 60 m in diameter!

Nuclear Notation

As you have learned, elements often have two or more isotopes. In other words, atoms of a given element always have the same number of protons, but may have differing numbers of neutrons.

How do scientists keep track of the number of protons and neutrons in an atom? The composition of an atom is often represented using just two numbers. The atomic number is the number of protons in the nucleus, which identifies the element. The mass number is the total number of protons and neutrons. Along with the atomic number, the mass number identifies a particular isotope of the element. Figure 1.21 shows how to use these numbers.

![Figure 1.20](image1.png)

**Figure 1.20** This illustration shows a simplified modern model of an atom. Notice that a fuzzy, cloud-like region surrounds the atomic nucleus. Electrons exist in this region at certain allowed energy levels.

![Figure 1.21](image2.png)

**Figure 1.21** Scientists represent isotopes using element symbols and by adding the mass number to an element name. These examples show how to represent isotopes of carbon and hydrogen. Hydrogen-2 and hydrogen-3 are more commonly known as deuterium and tritium.
You can determine the number of neutrons in a nucleus by subtracting the atomic number from the mass number.

\[
\text{number of neutrons} = \text{mass number} - \text{atomic number}
\]

For example, how many neutrons are in lithium-7?

- The number 7 in its name tells you that lithium-7 has a mass number of 7.
- Determine the atomic number of lithium by examining the periodic table in Appendix B.

Since lithium is the third element on the periodic table, its atomic number is 3. Therefore, the number of neutrons in lithium-7 is:

\[
\text{number of neutrons} = 7 - 3 = 4
\]

What about electrons? Atoms are electrically neutral, so they have no charge. Therefore, the number of negatively charged electrons must equal the number of positively charged protons. The atomic number of a neutral atom, then, is the same as the number of electrons in that atom. Answer the Practice Problems below to be sure you understand how to work with nuclear notation.

**Practice Problems**

5. State the number of neutrons in each of the following isotopes.
   (a) \( ^{22}_{10}\)Ne
   (b) \( ^{4}_{2}\)He
   (c) \( ^{40}_{20}\)Ca
   (d) \( ^{27}_{13}\)Al

6. State the number of each of the following subatomic particles in oxygen-17.
   (a) protons
   (b) electrons
   (c) neutrons

7. A certain isotope has a mass number of 35. The isotope has 18 neutrons.
   (a) What is the atomic number of the isotope?
   (b) What is the atomic symbol of the isotope?
   (c) Use nuclear notation to represent the isotope.

8. How many nucleons are in sodium-23?

**Section 1.2 Summary**

In this section, you examined some of the evidence that led to an atomic theory that involves protons and neutrons in a nucleus surrounded by electrons in energy levels. In the next section, you will see how the arrangement of these electrons affects how compounds form.
Check Your Understanding

1. Make a table that summarizes the following features of the electron, proton, and neutron: location within the atom, relative mass, and electrical charge.

2. Explain why Thomson’s theory of the atom could not explain the results of Rutherford’s gold foil experiment.

3. Explain how the following observations provided evidence for the existence of neutrons.
   (a) A helium atom has four times the mass of a hydrogen atom.
   (b) Different atoms of the same element can have different masses.

4. In your notebook, copy and complete the following table.

<table>
<thead>
<tr>
<th>Neutral isotopes</th>
<th>Number of neutrons</th>
<th>Number of protons</th>
<th>Number of electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>carbon-12</td>
<td>6</td>
<td>6</td>
<td>6</td>
</tr>
<tr>
<td>lithium-7</td>
<td>(a)</td>
<td>3</td>
<td>(b)</td>
</tr>
<tr>
<td>sodium-23</td>
<td>12</td>
<td>(c)</td>
<td>(d)</td>
</tr>
<tr>
<td>chlorine-37</td>
<td>(e)</td>
<td>(f)</td>
<td>(g)</td>
</tr>
</tbody>
</table>

5. Describe the reasoning that links each observation-hypothesis pair below.

<table>
<thead>
<tr>
<th>Scientist</th>
<th>Observation</th>
<th>Hypothesis</th>
</tr>
</thead>
<tbody>
<tr>
<td>Thomson</td>
<td>Cathode rays are attracted to positively charged plates. They curve with a measurable radius.</td>
<td>Atoms contain tiny, negatively charged particles. Atoms also contain sufficient positive charge to counteract the negative charge.</td>
</tr>
<tr>
<td>Rutherford</td>
<td>Most alpha particles pass through a gold foil, but a few rebound.</td>
<td>Atoms have a tiny, dense nucleus, but are mostly empty space.</td>
</tr>
<tr>
<td>Bohr</td>
<td>Hydrogen spectra consist of only a few, specific colours of light.</td>
<td>Electrons in atoms exist in specific energy levels.</td>
</tr>
</tbody>
</table>

6. **Thinking Critically** Give an example from the development of atomic theory that illustrates each of the following features of scientific thought.
   (a) A hypothesis allows you to make predictions that can be tested.
   (b) Theories must explain experimental observations.

7. **Thinking Critically** Rutherford studied for his undergraduate degree in New Zealand, did graduate studies to get a Ph.D. at Cambridge University in England, did further research at McGill University in Montréal, and then returned to England, to Manchester University. Do you think all this moving around was a disadvantage or an advantage to Rutherford as a scientist? Explain your answer.

8. **Apply** If an element contains two or more naturally occurring isotopes, is it a pure substance? Explain your answer.