Chapter 4, Lesson 1: Protons, Neutrons, and Electrons

Key Concepts

- Atoms are made of extremely tiny particles called protons, neutrons, and electrons.
- Protons and neutrons are in the center of the atom, making up the nucleus.
- Electrons surround the nucleus.
- Protons have a positive charge.
- Electrons have a negative charge.
- The charge on the proton and electron are exactly the same size but opposite.
- Neutrons have no charge.
- Since opposite charges attract, protons and electrons attract each other.

Summary

Students will put a static charge on a strip of plastic by pulling it between their fingers. They will see that the plastic is attrached to their fingers. Students will be introduced to the idea that rubbing the strip with their fingers caused electrons to move from their skin to the plastic giving the plastic a negative charge and their skin a positive charge. Through these activities, students will be introduced to some of the characteristics of electrons, protons, and neutrons, which make up atoms.

Objective

Students will be able to explain, in terms of electrons and protons, why a charged object is attracted or repelled by another charged object. They will also be able to explain why a charged object can even be attracted to an uncharged object. Students will also be able to explain that the attraction between positive protons and negative electrons holds an atom together.

Evaluation

The activity sheet will serve as the "Evaluate" component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding.

Safety

Be sure you and the students wear properly fitting goggles.

Materials for Each Group

- Plastic grocery bag
- Scissors
- Inflated balloon
- Small pieces of paper, confetti-size

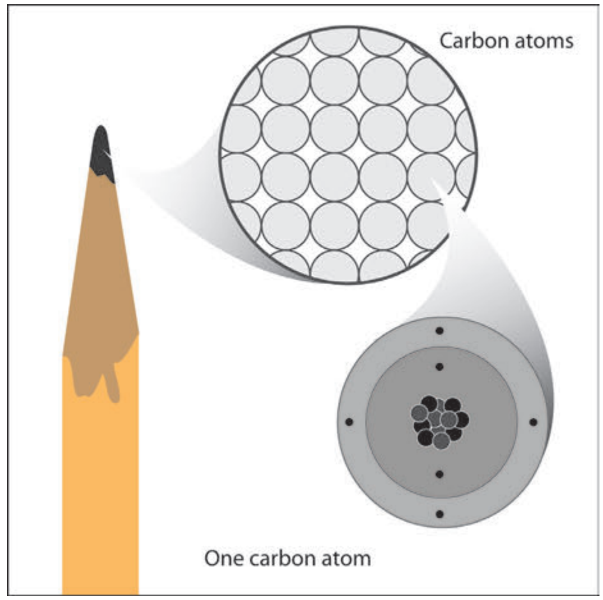
Materials for the Demonstration

- Sink
- Balloon

ENGAGE

1. Show a picture of a pencil point and how the carbon atoms look at the molecular level.

Project the image *Pencil Zoom.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson1#pencil_zoom



Students should be familiar with the parts of the atom from Chapter 3 but reviewing the main points is probably a good idea.

Ask students questions such as the following:

- **What are the three different tiny particles that make up an atom?**
  Protons, neutrons, and electrons.
- **Which of these is in the center of the atom?**
  Protons and neutrons are in the center (nucleus) of the atom. You may want to mention that hydrogen is the only atom that usually has no neutrons. The nucleus of most hydrogen atoms is composed of just 1 proton. A small percentage of hydrogen atoms have 1 or even 2 neutrons. Atoms of the same element with different numbers of neutrons are called *isotopes*. These will be discussed in Lesson 2.
- **What zooms around the nucleus of an atom?**
  Electrons
- **Which one has a positive charge, a negative charge, and no charge?**
  Proton—positive; electron—negative; neutron—no charge. The charge on the proton and electron are exactly the same size but opposite. The same number of protons and electrons exactly cancel one another in a neutral atom.

***Note**: The picture shows a simple model of the carbon atom. It illustrates some basic information like the number of protons and neutrons in the nucleus. It also shows that the number of electrons is the same as the number of protons. This model also shows that some electrons can be close to the nucleus and others are further away. One problem with this model is that it suggests that electrons orbit around the nucleus in perfect circles on the same plane, but this is not true. The more widely accepted model shows the electrons as a more 3-dimensional "electron cloud" surrounding the nucleus. Students will be introduced to these ideas in a bit more detail in Lesson 3. But for most of our study of chemistry at the middle school level, the model shown in the illustration will be very useful. Also, for most of our uses of this atom model, the nucleus will be shown as a dot in the center of the atom.*

2. Show animations and explain that protons and electrons have opposite charges and attract each other.

Project the animation *Protons and Electrons.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson1#protons_and_electrons

Explain to students that two protons repel each other and that two electrons repel each other. But a proton and an electron attract each other. Another way of saying this is that the same or "like" charges repel one another and opposite charges attract one another.

Since opposite charges attract each other, the negatively charged electrons are attracted to the positively charged protons. Tell students that this attraction is what holds the atom together.

Project the animation *Hydrogen Atom.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson1#hydrogen_atom

Explain to students that in a hydrogen atom, the negatively charged electron is attracted to the positively charged proton. This attraction is what holds the atom together.

Tell students that hydrogen is the simplest atom. It has only 1 proton, 1 electron, and 0 neutrons. It is the only atom that does not have any neutrons. Explain that this is a simple model that shows an electron going around the nucleus.

Click on the button "Show cloud" and explain to students that this is a different model. It shows the electron in the space surrounding the nucleus that is called an electron cloud or energy level. It is not possible to know the location of an electron but only the region where it is most likely to be. The electron cloud or energy level shows the region surrounding the nucleus where the electron is most likely to be.

Note*: Inquisitive students might ask how the positively charged protons are able to stay so close together in the nucleus: Why don't they repel each other? This is a great question. The answer is well beyond an introduction to chemistry for middle school, but one thing you can say is that there is a force called the "Strong Force," which holds protons and neutrons together in the nucleus of the atom. This force is much stronger than the force of repulsion of one proton from another.*

Another good question: Why doesn't the electron smash into the proton? If they are attracted to each other, why don't they just collide? Again, a detailed answer to this question is beyond the scope of middle school chemistry. But a simplified answer has to do with the energy or speed of the electron. As the electron gets closer to the nucleus, its energy and speed increases. It ends up *moving in a region surrounding the nucleus at a speed that is great enough to balance the attraction that is pulling it in, so the electron does not crash into the nucleus.*

Give each student an activity sheet.
Have students answer questions about the illustration on the activity sheet. Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions.

EXPLORE

3. Do an activity to show that electrons and protons attract each other.

Students can see evidence of the charges of protons and electrons by doing an activity with static electricity.

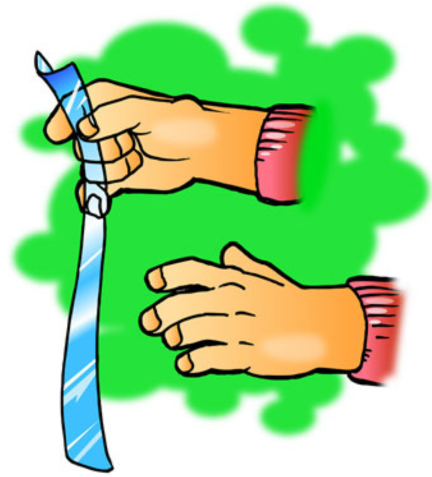
***Note**: When two materials are rubbed together in a static electricity activity, one material tends to lose electrons while the other material tends to gain electron. In this activity, human skin tends to lose electrons while the plastic bag, made of polyethylene, tends to gain electrons.*

Question to investigate
What makes objects attract or repel each other?

Materials for each group
- Plastic grocery bag
- Scissors

Procedure, part 1
Charged plastic and charged skin
1. Cut 2 strips from a plastic grocery bag so that each is about 2–4 cm wide and about 20 cm long.
2. Hold the plastic strip firmly at one end. Then grasp the plastic strip between the thumb and fingers of your other hand as shown.
3. Quickly pull your top hand up so that the plastic strip runs through your fingers. Do this three or four times.
4. Allow the strip to hang down. Then bring your other hand near it.

5. Write “attract” or “repel” in the chart on the activity sheet to describe what happened.



Expected results
The plastic will be attracted to your hand and move toward it. Students may notice that the plastic is also attracted to their arms and sleeves. Let students know that later in this lesson they will investigate why the plastic strip is also attracted to surfaces that have not been charged (neutral).

***Note**: If students find that their plastic strip does not move toward their hand, it must not have been charged well enough. Have them try charging their plastic strip by holding it down on their pants or shirt and then quickly pulling it with the other hand. Then they should test to see if the plastic is attracted to their clothes. If not, students should try charging the plastic again.*

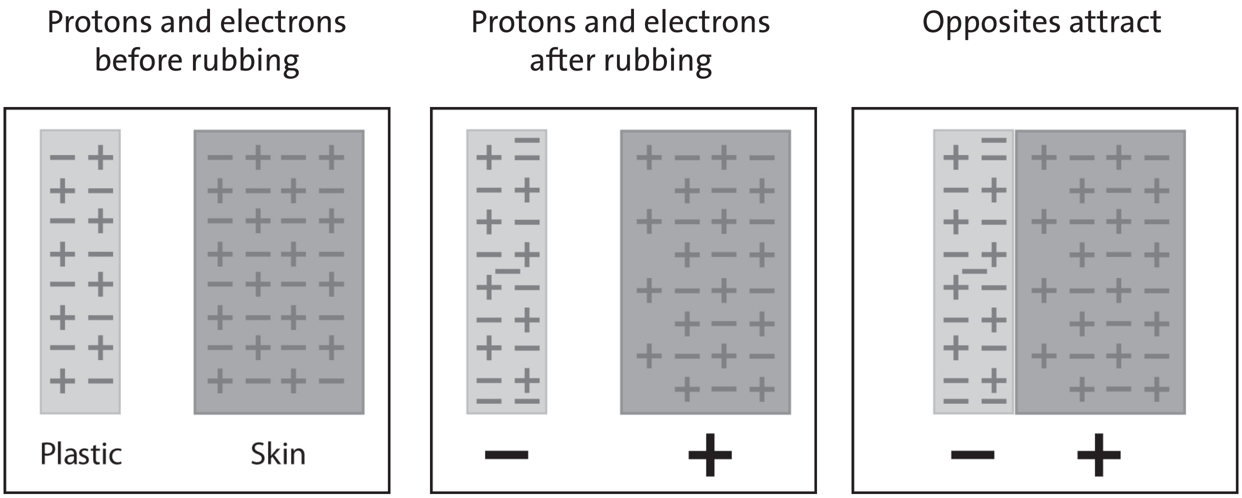
EXPLAIN

4. Show students models comparing the number of protons and electrons in the plastic and skin before and after rubbing them together.

Tell students that the plastic strip and their skin are made of molecules that are made of atoms. Tell students to assume that the plastic and their skin are neutral—that they have the same number of protons as electrons.

Project the image *Charged plastic and hand.*
www.middleschoolchemistry.com/multimedia/chapter 4/lesson1#charged_plastic_and_hand.jpg

Point out that before the students pulled the plastic between their fingers, the number of protons and electrons in each is the same. Then, when students pulled the plastic through their fingers, electrons from their skin got onto the plastic. Since the plastic has more electrons than protons, it has a negative charge. Since their fingers gave up some electrons, their skin now has more protons than electrons so it has a positive charge. The positive skin and the negative plastic attract each other because positive and negative attract.



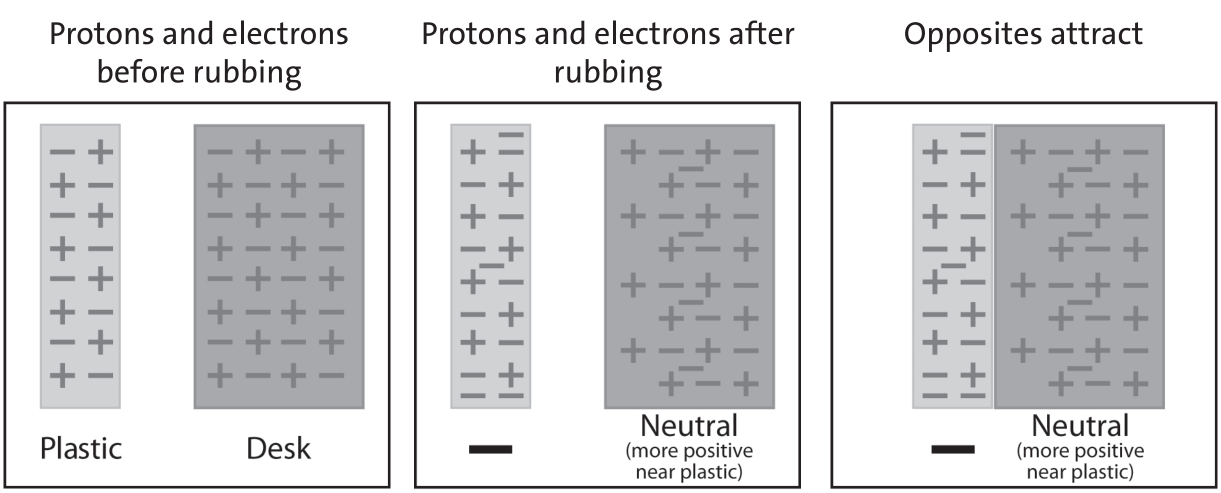
EXPLORE

5. Have students investigate what happens when a rubbed plastic strip is held near a desk or chair.

Procedure, part 2

Charged plastic and neutral desk

1. Charge one strip of plastic the same way you did previously.
2. This time, bring the plastic strip toward your desk or chair.
3. Write "attract" or "repel" in the chart.



Expected results
The plastic moves toward the desk.

Explain to students why the plastic is attracted to the desk. The answer takes a couple of steps, so you can guide students by drawing or projecting a magnified illustration of the plastic and desk.

After pulling the plastic between their fingers, the plastic gains extra electrons and a negative charge. The desk has the same number of protons as electrons and is neutral. When the plastic gets close to the desk, the negatively charged plastic repels electrons on the surface of the desk. This makes the surface of the desk near the plastic slightly positive. The negatively charged plastic is attracted to this positive area, so the plastic moves toward it.

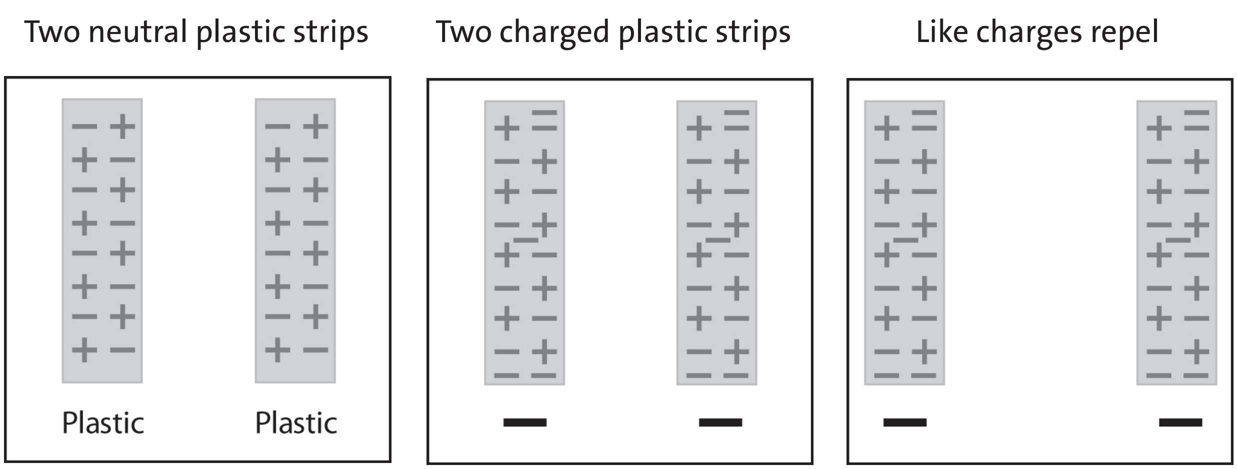
6. Have students charge two pieces of plastic and hold them near each other to see if electrons repel one other.

Ask students to make a prediction:

- What do you think will happen if you charge two strips of plastic and bring them near each other?

Procedure, part 3
2 pieces of charged plastic

1. Charge two strips of plastic
2. Slowly bring the two strips of plastic near each other.
3. Write "attract" or "repel" in the chart on the activity sheet.



Expected results
The strips will move away or repel each other. Since both strips have extra electrons on them, they each have extra negative charge. Since the same charges repel one another, the strips move away from each other.

Ask students:

- **What happened when you brought the two pieces of plastic near each other?**
  The ends of the strips moved away from each other.
- **Use what you know about electrons and charges to explain why this happens.**
  Each strip has extra electrons so they are both negatively charged. Because like charges repel, the pieces of plastic repelled each other.

EXTEND

7. Have students apply their understanding of protons and electrons to explain what happens when a charged balloon is brought near pieces of paper.

Materials for each group

- Inflated balloon
- Small pieces of paper, confetti-size

Procedure

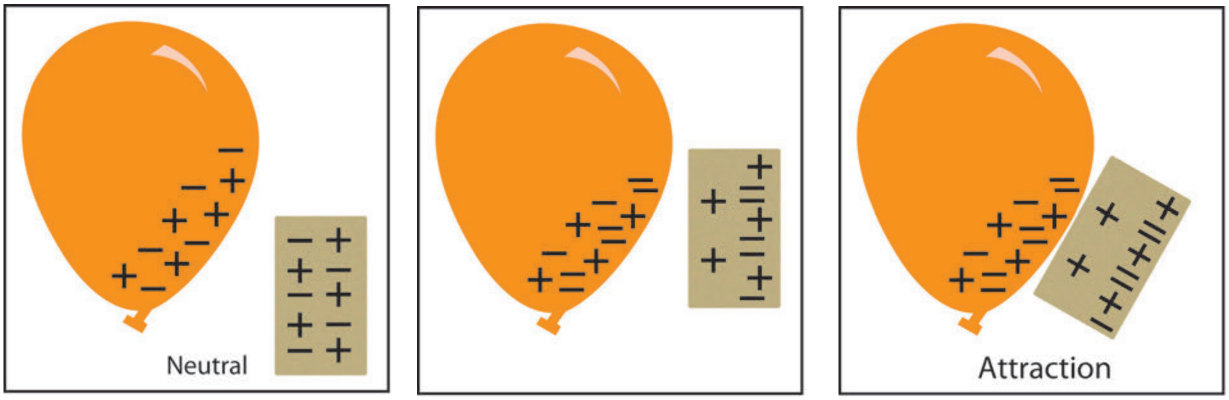
- Rub a balloon on your hair or clothes.
- Bring the balloon slowly toward small pieces of paper.

Expected results

The pieces of paper will jump up and stick on the balloon.

Ask students:

- **What did you observe when the charged balloon was held near the pieces of paper?**
  The paper pieces moved up and stuck on the balloon.
- **Use what you know about electrons, protons, and charges to explain why this happens.**
  When you rub the balloon on your hair or clothes it picks up extra electrons, giving the balloon a negative charge. When you bring the balloon near the paper, the electrons from the balloon repel the electrons in the paper. Since more protons are at the surface of the paper, it has a positive change. The electrons are still on the paper, just not at the surface, so overall the paper is neutral. Opposites attract, so the paper moves up toward the balloon.



Show the simulation *Balloons and Static Electricity* from the University of Colorado at Boulder's Physics Education Technology site.
http://phet.colorado.edu/simulations/sims.php?sim=Balloons_and_Static_Electricity

In the simulation, check the boxes "show all charges" and "wall". Uncheck everyting else. In this simulation, you can rub the balloon a little bit on the sweater and see that some of the electrons from the sweater move onto the balloon. This gives the balloon a negative charge. Since the sweater lost some electrons, it has more protons than electrons, so it has a positive charge. If you move the balloon toward the sweater, it will be attracted. This is like moving the charged plastic strip toward the cloth it was rubbed on.

You can also move the balloon toward the wall. The excess negative charge on the balloon repels negative charge on the surface of the wall. This leaves more positive charge on the surface of the wall. The negatively charged balloon is attracted to the positive area on the wall. This is like moving the charged plastic strip toward the finger.

EXTRA EXTEND

8. Demonstrate how electrons can attract a stream of water.

Either do the following demonstration or show the video *Balloon and Water*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson1#balloon_and_water

Materials for the demonstration
- Sink
- Balloon

Procedure
1. Rub a balloon on your shirt or pants to give it a static charge.
2. Turn on the faucet so that there is a very thin stream of water.
3. Slowly bring the charged part of the balloon close to the stream of water.

Expected results
The stream of water should bend as it is attracted to the balloon.

Ask students:

- **What did you observe when the charged balloon was held near the stream of water?**
  The stream of water bent toward the balloon.
- **Use what you know about electrons, protons, and charges to explain why this happens.**
  When you rub the balloon on your hair or clothes it picks up extra electrons, giving the balloon a negative charge. When you bring the balloon near the stream of water, the electrons from the balloon repel the electrons in the water. Since more protons are at the surface of the water, it has a positive change. Opposites attract, so the water moves toward the balloon.

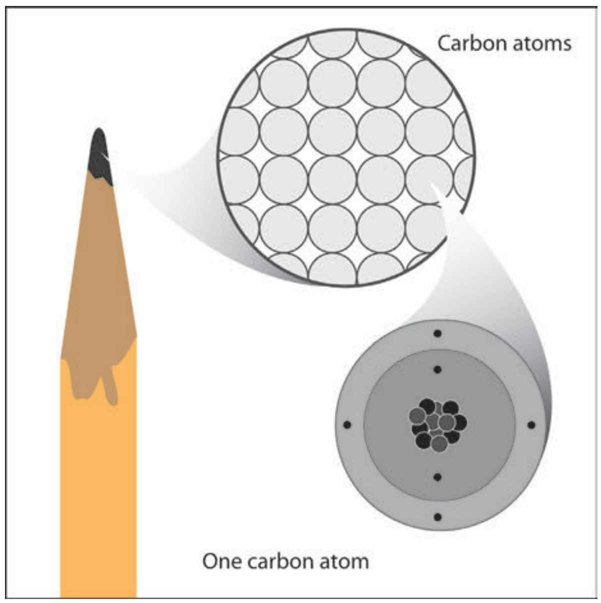
Activity Sheet
Chapter 4, Lesson 1
Protons, Neutrons, and Electrons

Name ____________________________

Date ____________________________

INTRODUCTION

If you look closely at the tip of a sharpened pencil, you will see that it is made of graphite. Going deeper, graphite is made of carbon atoms. Deeper still, each carbon atom is made of protons, neutrons, and electrons. In this lesson, you will explore these subatomic particles and their charges.



1. Label the nucleus (protons, neutrons) and electrons in the drawing of a carbon atom above.

2. Draw a line between the subatomic particle and its charge.

proton	no charge
electron	positive charge
neutron	negative charge

3. Would the following subatomic particles attract each other or repel one another?

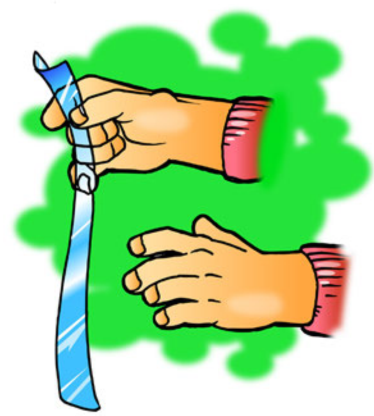
   Two protons _______________
   Two electrons ______________
   A proton and an electron __________

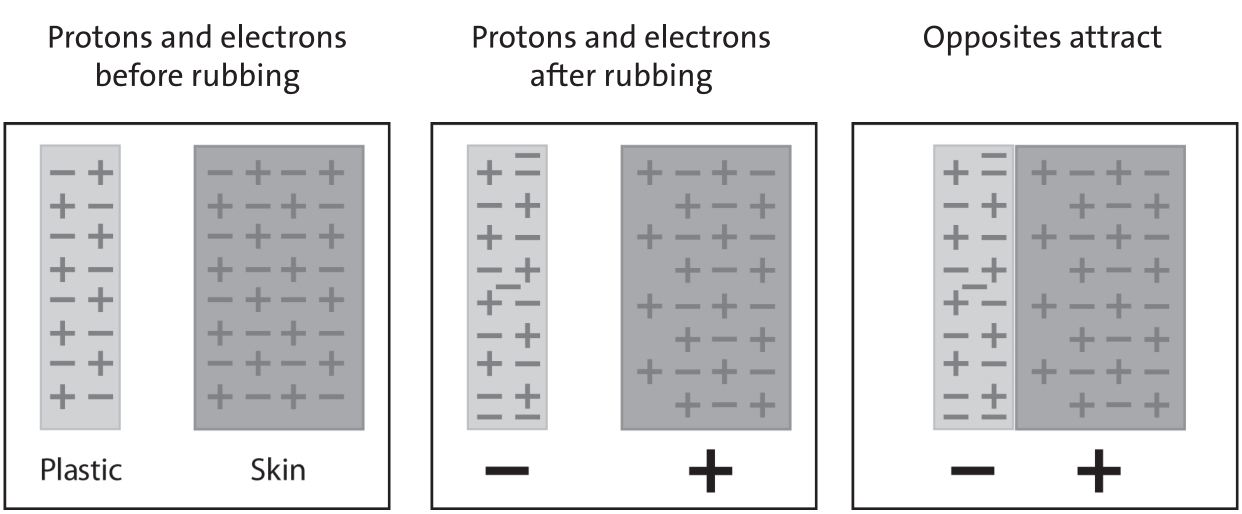
ACTIVITY

Question to investigate
What makes objects attract or repel each other?

Materials for each group
- Plastic grocery bag
- Scissors

Procedure, part 1
Charged plastic and charged skin
1. Cut 2 strips from a plastic grocery bag so that each is about 2–4 cm wide and about 20 cm long.
2. Hold the plastic strip firmly at one end. Then grasp the plastic strip between the thumb and fingers of your other hand as shown.
3. Quickly pull your top hand up so that the plastic strip runs through your fingers. Do this three or four times.
4. Allow the strip to hang down. Then bring your other hand near it.
5. Write "attract" or "repel" in the chart on page 256 to describe what happened.



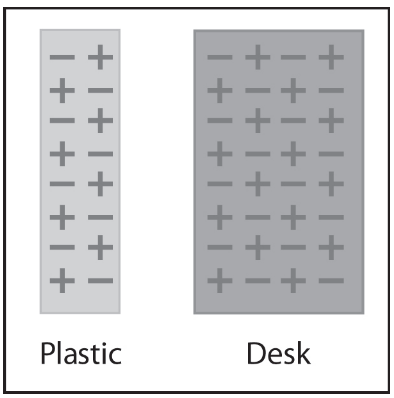


Procedure, part 2

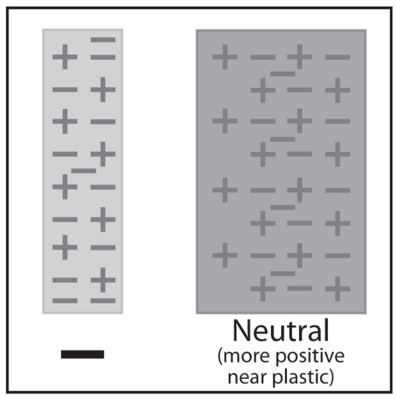
Charged plastic and neutral desk

1. Charge one strip of plastic the same way you did previously.
2. This time, bring the plastic strip toward your desk or chair.
3. Write “attract” or “repel” in the chart on the next page.

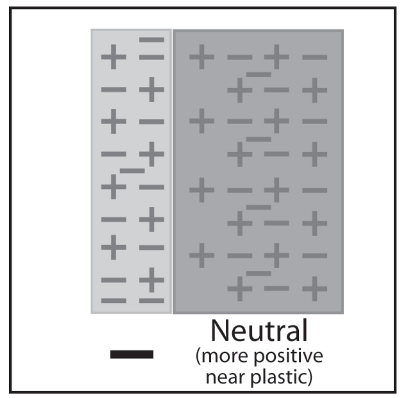
Protons and electrons before rubbing



Protons and electrons after rubbing



Opposites attract

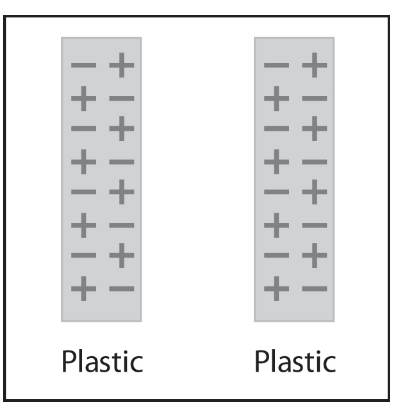


Procedure, part 3

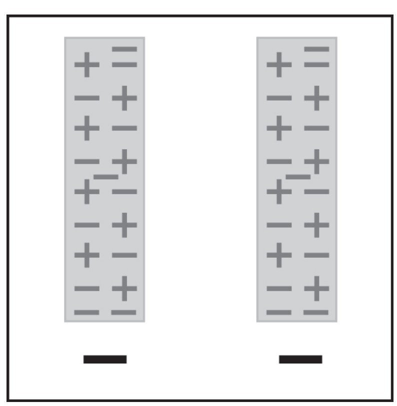
2 pieces of charged plastic

1. Charge two strips of plastic
2. Slowly bring the two strips of plastic near each other.
3. Write “attract” or “repel” in the chart on the next page.

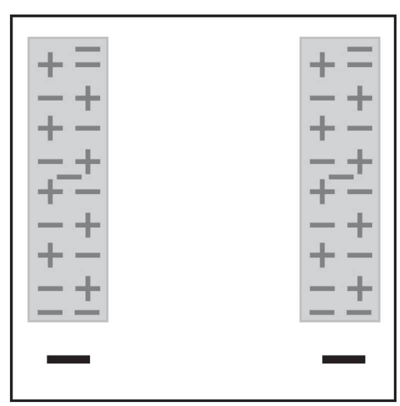
Two neutral plastic strips



Two charged plastic strips



Like charges repel



EXPLAIN IT WITH ATOMS & MOLECULES

What happened when you brought the following materials near each other?		
Materials	Attract or Repel	Use what you know about electrons, protons, and charges to explain your observations
charged plastic + charged skin		
charged plastic + neutral desk		
charged plastic + charged plastic		

TAKE IT FURTHER

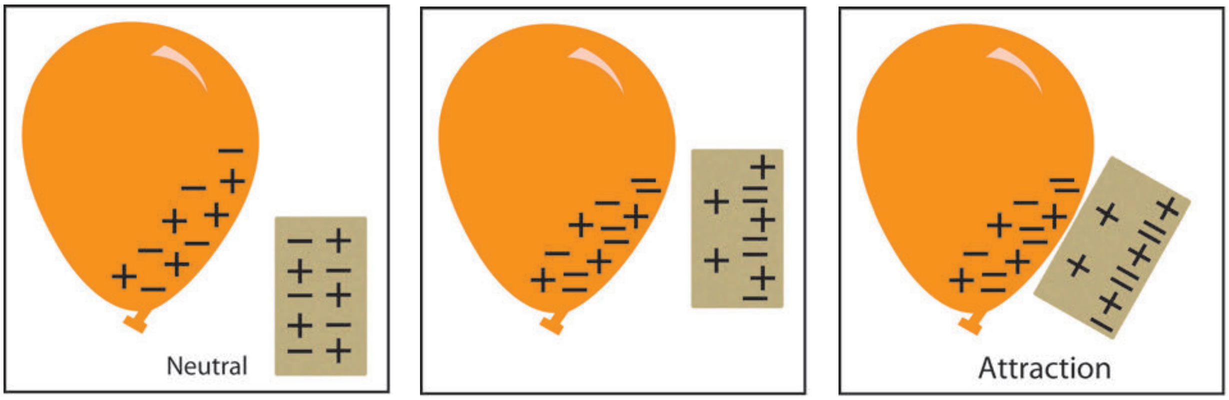
Materials for each group

- Inflated balloon
- Small pieces of paper, confetti-size

Procedure

- Rub a balloon on your hair or clothes.
- Bring the balloon slowly toward small pieces of paper.

4. Write captions beneath each picture explaining what happened between the balloon and your hair and the balloon and the paper in the activity.



Chapter 4, Lesson 2: The Periodic Table

Key Concepts

- The periodic table is a chart containing information about the atoms that make up all matter.
- An element is a substance made up of only one type of atom.
- The atomic number of an atom is equal to the number of protons in its nucleus.
- The number of electrons surrounding the nucleus of an atom is equal to the number of protons in its nucleus.
- Different atoms of the same element can have a different number of neutrons.
- Atoms of the same element with different numbers of neutrons are called "isotopes" of that element.
- The atomic mass of an element is the average mass of the different isotopes of the element.
- The atoms in the periodic table are arranged to show characteristics and relationships between atoms and groups of atoms.

Summary

Students will begin to look closely at the periodic table. They will be introduced to the basic information given for the elements in most periodic tables: the name, symbol, atomic number, and atomic mass for each element. Students will focus on the first 20 elements. They will try to correctly match cards with information about an element to each of the first 20 elements. Students will then watch several videos of some interesting chemical reactions involving some of these elements.

Objective

Students will identify different atoms by the number of protons in the nucleus and realize that the number of electrons equals the number of protons in a neutral atom. They will also be able to explain the meaning of atomic number and atomic mass.

Evaluation

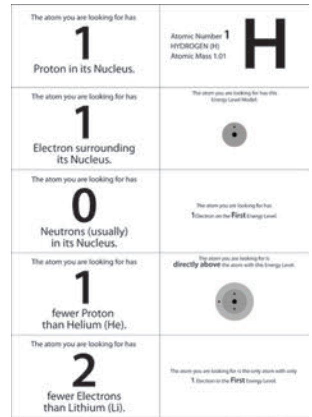
The activity sheet will serve as the "Evaluate" component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding.

About this Lesson

Lessons 2 and 3 both use the 20 atom description cards beginning on page 240.

Teacher preparation

Print out the 20 pages of element cards. The first page is shown. Laminate each page and cut out the cards. For Lesson 2, you will need the 5 cards for each element from the left side of each sheet. You will also need the card in the upper right corner. This is the atom name card. Tape each of the 20 atom name cards to a spot in the room where students can place the cards that match that atom nearby. For Lesson 3, you will need the atom name card, taped in the same location in the room, and the four cards beneath it. Divide the class into 10 groups of 2 or 3 students each.



ENGAGE

1. Introduce students to the periodic table.

Project the image *Periodic Table.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson2#periodic_table

Tell students that this is the periodic table. Explain that each box contains information about a different atom. The periodic table shows all the atoms that everything in the known universe is made from. It's kind of like the alphabet in which only 26 letters, in different combinations, make up many thousands of words. The 100 or so atoms of the periodic table, in different combinations, make up millions of different substances.

The Periodic Table of the Elements

3 — Atomic Number
Li — Element Symbol
Lithium — Element Name
6.94 — Average Atomic Mass

1 H Hydrogen 1.01																	2 He Helium 4.00
3 Li Lithium 6.94	4 Be Beryllium 9.01											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.39	31 Ga Gallium 69.72	32 Ge Germanium 72.61	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybednum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.91	56 Ba Barium 137.33	57 La Lanthanum 138.91	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.84	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.98	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium 178.49	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (269)	109 Mt Meitnerium (268)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (272)	112 Cn Copernicium (285)						

58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.97
90 Th Thorium 232.04	91 Pa Proactinium 231.04	92 U Uranium 238.03	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium 168.93	102 No Nobelium (259)	103 Lr Lawrencium (262)

Note: It is often confusing for students to see the terms "atom" and "element" used interchangeably as if they are the same thing. Explain to students that an atom is the smallest particle or "building block" of a substance. An element is a substance made up of all the same type of atom. For instance, a piece of pure carbon is made up of only carbon atoms. This piece of pure carbon is a sample of the element carbon. The people who developed the periodic table could have called it the Periodic Table of the Atoms but they did not have a firm understanding of atoms at that time. Since they were working with actual samples of elements such as copper, mercury, sulfur, etc., they called it the periodic table of the elements.

Optional
Play one or both of the following songs.

- *The Elements* by Tom Lehrer with animation by Mike Stanfill
  www.privatehand.com/flash/elements.html
- *Meet the Elements* by They Might be Giants
  www.youtube.com/watch?v=d0zION8xjbM

2. Explain the meaning of the numbers and letters in the boxes in the periodic table.

Tell students that the class will focus on the first 20 elements over 2 days. On the first day, they will look at the number of protons, electrons, and neutrons in the atoms of each element. On the second day, they will look at the arrangement of electrons in the atoms.

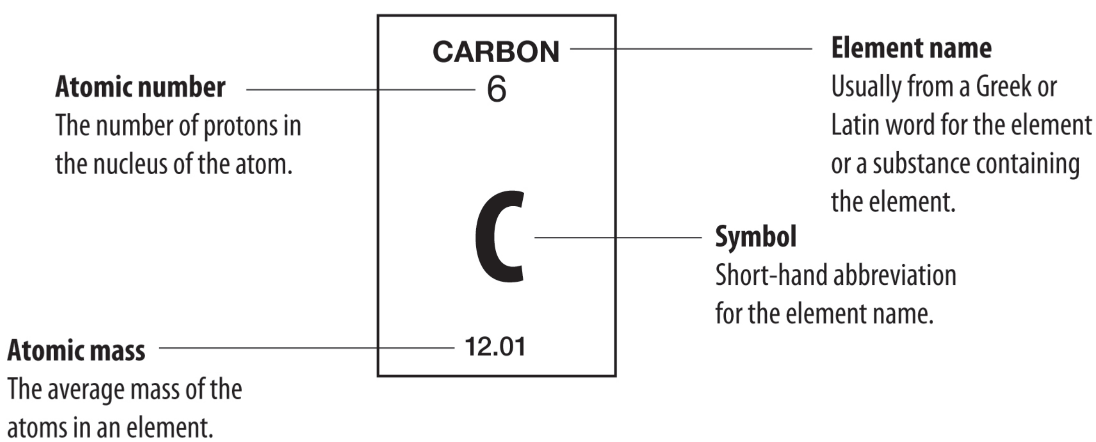
Give each student a copy of the periodic table of the elements, the periodic table of elements 1–20, and the activity sheet.
Students will use the periodic table of elements 1–20, along with the activity sheet, in the lesson they will do today.

Project the image *Periodic Table of the First 20 Elements*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson2#first_twenty

PERIODIC TABLE
ELEMENTS 1–20

HYDROGEN 1 **H** 1.01							HELIUM 2 **He** 4.00
LITHIUM 3 **Li** 6.94	BERYLLIUM 4 **Be** 9.01	BORON 5 **B** 10.81	CARBON 6 **C** 12.01	NITROGEN 7 **N** 14.01	OXYGEN 8 **O** 16.00	FLUORINE 9 **F** 19.00	NEON 10 **Ne** 20.18
SODIUM 11 **Na** 22.99	MAGNESIUM 12 **Mg** 24.31	ALUMINUM 13 **Al** 26.98	SILICON 14 **Si** 28.09	PHOSPHORUS 15 **P** 30.97	SULFUR 16 **S** 32.07	CHLORINE 17 **Cl** 35.45	ARGON 18 **Ar** 39.95
POTASSIUM 19 **K** 39.10	CALCIUM 20 **Ca** 40.08						

Project the image *Element explanation*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson2#element_explanation
Explain what the numbers and letters in each box on the periodic table represent.



Explain atomic mass.
The atomic mass of an element is based on the mass of the protons, neutrons, and electrons of the atoms of that element. The mass of the proton and neutron are about the same, but the mass of the electron is much smaller (about 1/2000 the mass of the proton or neutron). The majority of the atomic mass is contributed by the protons and neutrons.

For any element in the periodic table, the number of electrons in an atom of that element always equals the number of protons in the nucleus. But this is not true for neutrons. Atoms of the same element can have different numbers of neutrons than protons. Atoms of the same element with different numbers of neutrons are called *isotopes* of that element. The atomic mass in the periodic table is an average of the atomic mass of the isotopes of an element. For the atoms of the first 20 elements, the number of neutrons is either equal to or slightly greater than the number of protons.

For example, the vast majority of carbon atoms have 6 protons and 6 neutrons, but a small percentage have 6 protons and 7 neutrons, and an even smaller percentage have 6 protons and 8 neutrons. Since the majority of carbon atoms have a mass very close to 12, and only a small percentage are greater than 12, the average atomic mass is slightly greater than 12.

3. Describe the activity students will do to learn about the first 20 elements of the periodic table.

Show students that you have 100 cards (5 for each of the first 20 elements). Explain that each card contains information about one of the first 20 atoms of the periodic table. The students' job is to read the card carefully, figure out which atom the card is describing, and put the card at the spot in the room for that atom.

Review the information about protons, electrons, and neutrons students need to know in order to match the cards with the correct element:

Proton
- Positively charged particle in the nucleus of the atom.
- The number of protons in an atom's nucleus is the atomic number.

Electron
- Negatively charged particle surrounding the nucleus of the atom.
- The number of electrons surrounding the nucleus of an atom is equal to the number of protons in the atom's nucleus.

Neutron

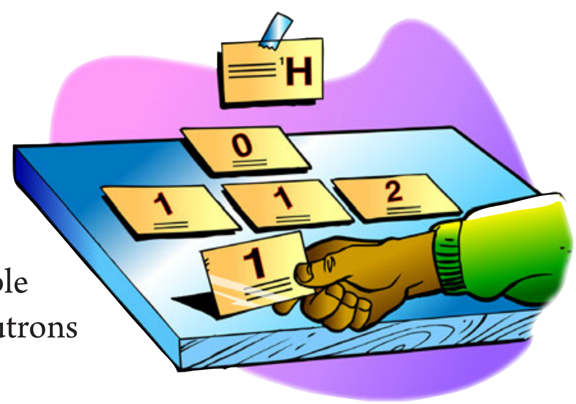
- Particle in the nucleus that has almost the same mass as a proton but has no charge.
- For the atoms of the first 20 elements, the number of neutrons is either equal to or slightly greater than the number of protons.

To match the number of neutrons listed on your card to the correct element, look for an element on the periodic table so that if you add the number of neutrons on your card to the protons of the element, you will get close to the atomic mass for that element.
For example, you may have a card that says that the atom you are looking for has 5 neutrons. You would look at the periodic table to find an atom that you could add 5 to its number of protons that would give you a sum close to the atomic mass given for that element. The answer is beryllium (Be), which has 4 protons and an atomic mass of 9.01.

Note: There are a few neutron cards that have two possible correct elements instead of just one:

- 6 Neutrons—Boron or Carbon
- 10 Neutrons—Fluorine or Neon
- 12 Neutrons—Sodium or Magnesium
- 16 Neutrons—Phosphorous or Sulfur
- 20 Neutrons—Potassium or Calcium

EXPLORE



4. Have groups work together to place each card with its correct atom.

Distribute the cards to groups. If you have 10 groups, each group will get 10 cards. Be available to help students who have trouble with the neutrons and atomic mass.

5. Discuss the placement of the cards for two or three atoms.

Select two or three atoms and review whether the cards were placed correctly. This review will help reinforce the concepts about the structure of atoms and help students determine the number of protons, electrons, and neutrons in each type of atom.

Have students begin filling out the activity sheet with the following information:

- Number of protons
- Number of electrons
- Number of neutrons (usually)

PERIODIC TABLE ELEMENTS 1–20

Write the number of protons, electrons, and neutrons in each element.

HYDROGEN 1 # of Protons: 1 # of Electrons: 1 # of Neutrons: 0 **1.01**							**HELIUM** 2 # of Protons: 2 # of Electrons: 2 # of Neutrons: 2 **4.00**
LITHIUM 3 # of Protons: 3 # of Electrons: 3 # of Neutrons: 4 **6.94**	**BERYLLIUM** 4 # of Protons: 4 # of Electrons: 4 # of Neutrons: 5 **9.01**	**BORON** 5 # of Protons: 5 # of Electrons: 5 # of Neutrons: 6 **10.81**	**CARBON** 6 # of Protons: 6 # of Electrons: 6 # of Neutrons: 6 **12.01**	**NITROGEN** 7 # of Protons: 7 # of Electrons: 7 # of Neutrons: 7 **14.01**	**OXYGEN** 8 # of Protons: 8 # of Electrons: 8 # of Neutrons: 8 **16.00**	**FLUORINE** 9 # of Protons: 9 # of Electrons: 9 # of Neutrons: 10 **19.00**	**NEON** 10 # of Protons: 10 # of Electrons: 10 # of Neutrons: 10 **20.18**
SODIUM 11 # of Protons: 11 # of Electrons: 11 # of Neutrons: 12 **22.99**	**MAGNESIUM** 12 # of Protons: 12 # of Electrons: 12 # of Neutrons: 12 **24.31**	**ALUMINUM** 13 # of Protons: 13 # of Electrons: 13 # of Neutrons: 14 **26.98**	**SILICON** 14 # of Protons: 14 # of Electrons: 14 # of Neutrons: 14 **28.09**	**PHOSPHORUS** 15 # of Protons: 15 # of Electrons: 15 # of Neutrons: 16 **30.97**	**SULFUR** 16 # of Protons: 16 # of Electrons: 16 # of Neutrons: 16 **32.07**	**CHLORINE** 17 # of Protons: 17 # of Electrons: 17 # of Neutrons: 18 **35.45**	**ARGON** 18 # of Protons: 18 # of Electrons: 18 # of Neutrons: 22 **39.95**
POTASSIUM 19 # of Protons: 19 # of Electrons: 19 # of Neutrons: 20 **39.10**	**CALCIUM** 20 # of Protons: 20 # of Electrons: 20 # of Neutrons: 20 **40.08**						

Note: The number of neutrons may be different in the atoms of the same element. The atoms of an element with different numbers of nuetrons are called isotopes of that element. The number of neutrons shown in the chart represents the most common isotope for that element.

EXTEND

6. Introduce students to their element project and an online resource that they can use.

Assign each student to an element. Include the first 20 elements and any other elements that you find interesting so that each student can research and present their own.

Each student should find and present some basic information about their element to the class. The presentation can be in the form of a poster, pamphlet, PowerPoint presentation or other form. The presentations should be short and can include: atom name, atomic number, derivation of name, when and where discovered, natural sources of the element, major uses, and any other information you find important.

Some Internet sources for this information can be overwhelming. They can also contain advertising that you may not want students exploring. For basic information about the periodic table, including some images and video, *The Journal of Chemical Education's Periodic Table Live* is an excellent resource.

www.chemeddl.org/collections/ptl/

If there is time available, have students work on this atom project during the week.

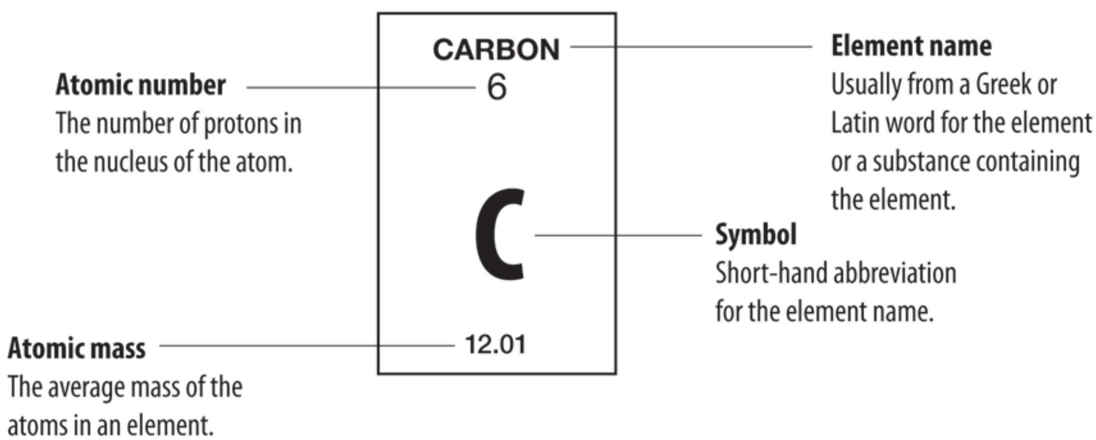
Activity Sheet
Chapter 4, Lesson 2
The Periodic Table

Name ____________________

Date ____________________

Your group will receive a set of cards with information that describes a particular atom. Your job is to figure out which atom the card describes and to place it in the area in your classroom for that atom.

You will use the Periodic Table, Elements 1–20 chart to help you determine what atom your card describes. The diagram and information below will help you match your cards to the correct atoms.



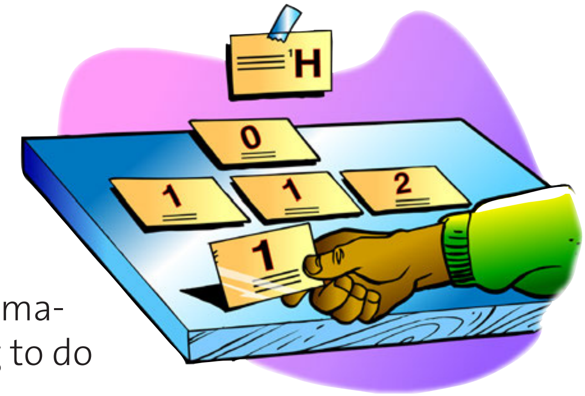
Parts of an Atom

Proton
Positively charged particle in the nucleus of the atom.
The number of protons in an atom's nucleus is the atomic number.

Electron
Negatively charged particle surrounding the nucleus of the atom.
The number of electrons surrounding the nucleus of an atom is equal to the number of protons in the atom's nucleus.

Neutron
Particle in the nucleus that has about the same mass as a proton but has no charge. For the atoms of the first 20 elements, the number of neutrons is either equal to or slightly greater than the number of protons.

Placing your cards



Once you know what the information in each box on your periodic table stands for and you know the parts of the atom, you will be able to correctly place most of your cards with the atoms they describe. You will need to know the following additional information in order to answer any question having to do with neutrons.

To match the number of neutrons listed on your card to the correct element, look for an element on the periodic table so that if you add the number of neutrons on your card to the protons of the element, you will get close to the atomic mass for that element.

For example, you may have a card that says, "The atom you are looking for has 5 neutrons." Look at the periodic table to find an atom that you could add 5 to its number of protons that would give you a sum close to the atomic mass given for that element. The answer is beryllium (Be), which has 4 protons and an atomic mass of 9.01.

PERIODIC TABLE
ELEMENTS 1–20

Write the number of protons, electrons, and neutrons in each element.

HYDROGEN 1 # of Protons: # of Electrons: # of Neutrons: **1.01**							**HELIUM** 2 # of Protons: # of Electrons: # of Neutrons: **4.00**
LITHIUM 3 # of Protons: # of Electrons: # of Neutrons: **6.94**	**BERYLLIUM** 4 # of Protons: # of Electrons: # of Neutrons: **9.01**	**BORON** 5 # of Protons: # of Electrons: # of Neutrons: **10.81**	**CARBON** 6 # of Protons: # of Electrons: # of Neutrons: **12.01**	**NITROGEN** 7 # of Protons: # of Electrons: # of Neutrons: **14.01**	**OXYGEN** 8 # of Protons: # of Electrons: # of Neutrons: **16.00**	**FLUORINE** 9 # of Protons: # of Electrons: # of Neutrons: **19.00**	**NEON** 10 # of Protons: # of Electrons: # of Neutrons: **20.18**
SODIUM 11 # of Protons: # of Electrons: # of Neutrons: **22.99**	**MAGNESIUM** 12 # of Protons: # of Electrons: # of Neutrons: **24.31**	**ALUMINUM** 13 # of Protons: # of Electrons: # of Neutrons: **26.98**	**SILICON** 14 # of Protons: # of Electrons: # of Neutrons: **28.09**	**PHOSPHORUS** 15 # of Protons: # of Electrons: # of Neutrons: **30.97**	**SULFUR** 16 # of Protons: # of Electrons: # of Neutrons: **32.07**	**CHLORINE** 17 # of Protons: # of Electrons: # of Neutrons: **35.45**	**ARGON** 18 # of Protons: # of Electrons: # of Neutrons: **39.95**
POTASSIUM 19 # of Protons: # of Electrons: # of Neutrons: **39.10**	**CALCIUM** 20 # of Protons: # of Electrons: # of Neutrons: **40.08**						

Note: Remember that the number of neutrons is not the same for every atom of an element. The number of neutrons you write in this chart will be a number, that when added to the number of protons, gives a sum as close as possible to the atomic mass.

The Periodic Table of the Elements

3 — Atomic Number
Li — Element Symbol
Lithium — Element Name
6.94 — Average Atomic Mass

1 **H** Hydrogen 1.01																	2 **He** Helium 4.00
3 **Li** Lithium 6.94	4 **Be** Beryllium 9.01											5 **B** Boron 10.81	6 **C** Carbon 12.01	7 **N** Nitrogen 14.01	8 **O** Oxygen 16.00	9 **F** Fluorine 19.00	10 **Ne** Neon 20.18
11 **Na** Sodium 22.99	12 **Mg** Magnesium 24.31											13 **Al** Aluminum 26.98	14 **Si** Silicon 28.09	15 **P** Phosphorus 30.97	16 **S** Sulfur 32.07	17 **Cl** Chlorine 35.45	18 **Ar** Argon 39.95
19 **K** Potassium 39.10	20 **Ca** Calcium 40.08	21 **Sc** Scandium 44.96	22 **Ti** Titanium 47.87	23 **V** Vanadium 50.94	24 **Cr** Chromium 52.00	25 **Mn** Manganese 54.94	26 **Fe** Iron 55.85	27 **Co** Cobalt 58.93	28 **Ni** Nickel 58.69	29 **Cu** Copper 63.55	30 **Zn** Zinc 65.39	31 **Ga** Gallium 69.72	32 **Ge** Germanium 72.61	33 **As** Arsenic 74.92	34 **Se** Selenium 78.96	35 **Br** Bromine 79.90	36 **Kr** Krypton 83.80
37 **Rb** Rubidium 85.47	38 **Sr** Strontium 87.62	39 **Y** Yttrium 88.91	40 **Zr** Zirconium 91.22	41 **Nb** Niobium 92.91	42 **Mo** Molybednum 95.94	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.07	45 **Rh** Rhodium 102.91	46 **Pd** Palladium 106.42	47 **Ag** Silver 107.87	48 **Cd** Cadmium 112.41	49 **In** Indium 114.82	50 **Sn** Tin 118.71	51 **Sb** Antimony 121.76	52 **Te** Tellurium 127.60	53 **I** Iodine 126.90	54 **Xe** Xenon 131.29
55 **Cs** Cesium 132.91	56 **Ba** Barium 137.33	57 **La** Lanthanum 138.91	72 **Hf** Hafnium 178.49	73 **Ta** Tantalum 180.95	74 **W** Tungsten 183.84	75 **Re** Rhenium 186.21	76 **Os** Osmium 190.23	77 **Ir** Iridium 192.22	78 **Pt** Platinum 195.08	79 **Au** Gold 196.97	80 **Hg** Mercury 200.59	81 **Tl** Thallium 204.38	82 **Pb** Lead 207.2	83 **Bi** Bismuth 208.98	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
87 **Fr** Francium (223)	88 **Ra** Radium (226)	89 **Ac** Actinium (227)	104 **Rf** Rutherfordium 178.49	105 **Db** Dubnium (262)	106 **Sg** Seaborgium (266)	107 **Bh** Bohrium (264)	108 **Hs** Hassium (269)	109 **Mt** Meitnerium (268)	110 **Ds** Darmstadtium (281)	111 **Rg** Roentgenium (272)	112 **Cn** Copernicium (285)						

58 **Ce** Cerium 140.12	59 **Pr** Praseodymium 140.91	60 **Nd** Neodymium 144.24	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.36	63 **Eu** Europium 151.96	64 **Gd** Gadolinium 157.25	65 **Tb** Terbium 158.93	66 **Dy** Dysprosium 162.50	67 **Ho** Holmium 164.93	68 **Er** Erbium 167.26	69 **Tm** Thulium 168.93	70 **Yb** Ytterbium 173.04	71 **Lu** Lutetium 174.97
90 **Th** Thorium 232.04	91 **Pa** Proactinium 231.04	92 **U** Uranium 238.03	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einstenium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium 168.93	102 **No** Nobelium (259)	103 **Lr** Lawrencium (262)

The atom you are looking for has

1

Proton in its Nucleus.

Atomic Number **1**
HYDROGEN (H)
Atomic Mass 1.01

H

The atom you are looking for has

1

Electron surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:

The atom you are looking for has

0

Neutrons (usually) in its Nucleus.

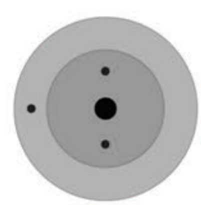
The atom you are looking for has
1 Electron on the **First** Energy Level.

The atom you are looking for has

1

fewer Proton than Helium (He).

The atom you are looking for is **directly above** the atom with this Energy Level.



The atom you are looking for has

2

fewer Electrons than Lithium (Li).

The atom you are looking for is the only atom with only
1 Electron in the **First** Energy Level.

The atom you are looking for has

2

Protons in its Nucleus.

Atomic Number **2**
HELIUM (He)
Atomic Mass 4.00

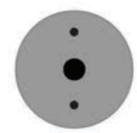
He

The atom you are looking for has

2

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

2

Neutrons (usually) in its Nucleus.

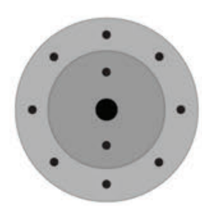
The atom you are looking for has **2** Electrons on the **First** Energy Level and no other electrons.

The atom you are looking for has

1

more Proton than Hydrogen (H).

The atom you are looking for is **directly above** the atom with this Energy Level.

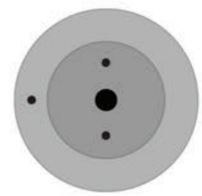
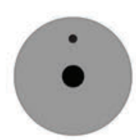
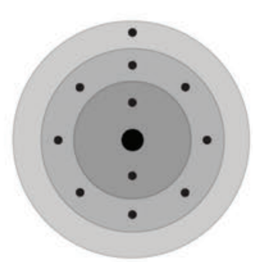


The atom you are looking for has

2

fewer Electrons than Beryllium (Be).

The atom you are looking for is the only atom with only **2** Electrons in the First Energy Level and no other electrons on any other level.

The atom you are looking for has **3** Protons in its Nucleus.	Atomic Number **3** LITHIUM (Li) Atomic Mass 6.94 **Li**
The atom you are looking for has **3** Electrons surrounding its Nucleus.	The atom you are looking for has this Energy Level Model: 
The atom you are looking for has **4** Neutrons (usually) in its Nucleus.	The atom you are looking for has **2** Electrons on the **First** Energy Level and **1** Electron on the **Second** Energy Level.
The atom you are looking for has **3** fewer Protons than Carbon (C).	The atom you are looking for is **directly below** the atom with this Energy Level. 
The atom you are looking for has **2** fewer Electrons than Boron.	The atom you are looking for is **directly above** the atom with this Energy Level. 

The atom you are looking for has

4

Protons in its Nucleus.

Atomic Number **4**
BERYLLIUM (Be)
Atomic Mass 9.01

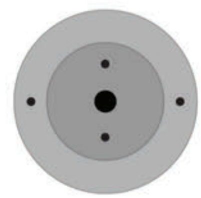
Be

The atom you are looking for has

4

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

5

Neutrons (usually) in its Nucleus.

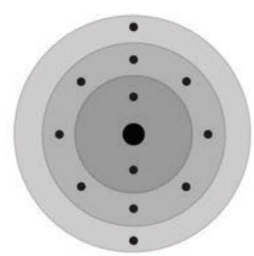
The atom you are looking for has
2 Electrons on the **First** Energy Level and
2 Electrons on the **Second** Energy Level.

The atom you are looking for has

4

fewer Protons than Oxygen (O).

The atom you are looking for is **directly above** the atom with this Energy Level.

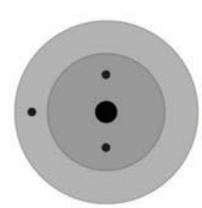


The atom you are looking for has

3

fewer Electrons than Nitrogen (N).

The atom you are looking for is **directly to the right** of the atom with this Energy Level.



The atom you are looking for has

5

Protons in its Nucleus.

Atomic Number **5**
Boron (B)
Atomic Mass 10.81

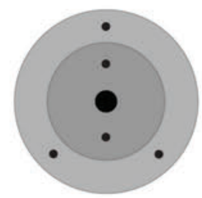
B

The atom you are looking for has

5

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

6

Neutrons (usually) in its Nucleus.

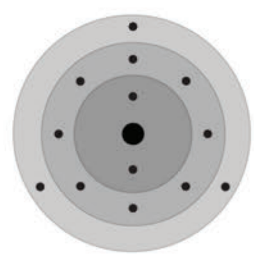
The atom you are looking for has
2 Electrons on the **First** Energy Level and
3 Electrons on the **Second** Energy Level.

The atom you are looking for has

4

more Protons than Hydrogen (H).

The atom you are looking for is **directly above** the atom with this Energy Level.

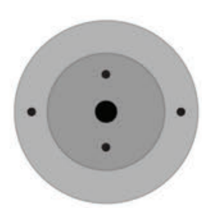


The atom you are looking for has

3

more Electrons than Helium (He).

The atom you are looking for is **directly to the right** of the atom with this Energy Level.



The atom you are looking for has

6

Protons in its Nucleus.

Atomic Number **6**
Carbon (C)
Atomic Mass 12.01

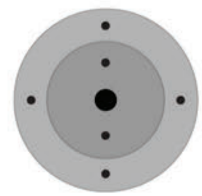
C

The atom you are looking for has

6

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

6

Neutrons (usually) in its Nucleus.

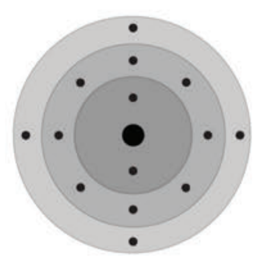
The atom you are looking for has
2 Electrons on the **First** Energy Level and
4 Electrons on the **Second** Energy Level.

The atom you are looking for has

1

more Proton than Boron (B).

The atom you are looking for is **directly above** the atom with this Energy Level.

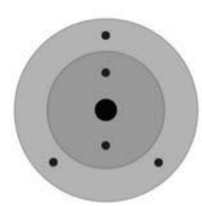


The atom you are looking for has

3

more Electrons than Lithium (Li).

The atom you are looking for is **directly to the right** of the atom with this Energy Level.



The atom you are looking for has

7

Protons in its Nucleus.

Atomic Number **7**
Nitrogen (N)
Atomic Mass 14.01

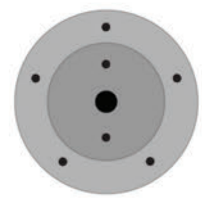
N

The atom you are looking for has

7

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

7

Neutrons (usually) in its Nucleus.

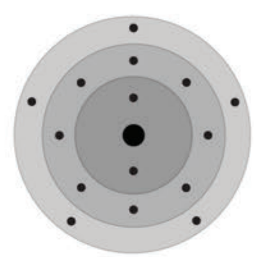
The atom you are looking for has
2 Electrons on the **First** Energy Level and
5 Electrons on the **Second** Energy Level.

The atom you are looking for has

3

fewer Protons than Neon (Ne).

The atom you are looking for is
directly above the atom with this Energy Level.

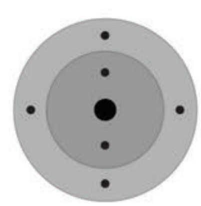


The atom you are looking for has

1

fewer Electron than Oxygen (O).

The atom you are looking for is
directly to the right
of the atom with this Energy Level.



The atom you are looking for has

8

Protons in its Nucleus.

Atomic Number **8**
Oxygen (O)
Atomic Mass 16.00

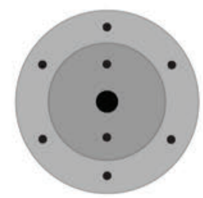
O

The atom you are looking for has

8

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

8

Neutrons (usually)
in its Nucleus.

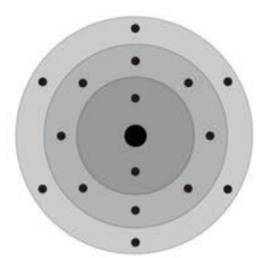
The atom you are looking for has
2 Electrons on the **First** Energy Level and
6 Electrons on the **Second** Energy Level.

The atom you are looking for has

2

more Protons
than Carbon (C).

The atom you are looking for is
directly above the atom with this Energy Level.

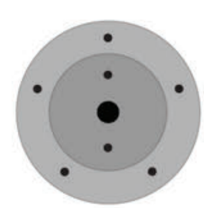


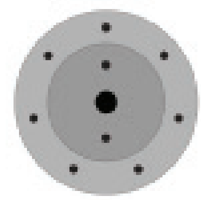
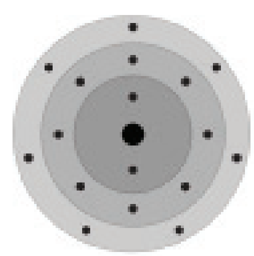
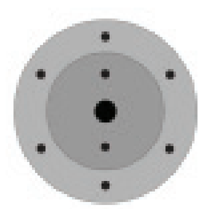
The atom you are looking for has

6

more Electrons
than Helium (He).

The atom you are looking for is
directly to the right
of the atom with this Energy Level.



The atom you are looking for has **9** Protons in its Nucleus.	Atomic Number **9** Fluorine (F) Atomic Mass 18.99 **F**
The atom you are looking for has **9** Electrons surrounding its Nucleus.	The atom you are looking for has this Energy Level Model: 
The atom you are looking for has **10** Neutrons (usually) in its Nucleus.	The atom you are looking for has **2** Electrons on the **First** Energy Level and **7** Electrons on the **Second** Energy Level.
The atom you are looking for has **1** fewer Proton than Neon (Ne).	The atom you are looking for is **directly above** the atom with this Energy Level. 
The atom you are looking for has **2** more Electrons than Nitrogen (N).	The atom you are looking for is **directly to the right** of the atom with this Energy Level. 

The atom you are looking for has

10

Protons in its Nucleus.

Atomic Number **10**
Neon (Ne)
Atomic Mass 20.18

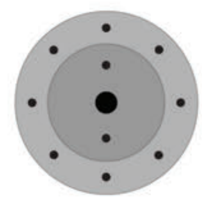
Ne

The atom you are looking for has

10

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

10

Neutrons (usually) in its Nucleus.

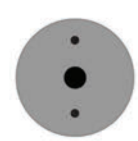
The atom you are looking for has
2 Electrons on the **First** Energy Level and
8 Electrons on the **Second** Energy Level.

The atom you are looking for has

8

more Protons than Helium (He).

The atom you are looking for is **directly below** the atom with this Energy Level.

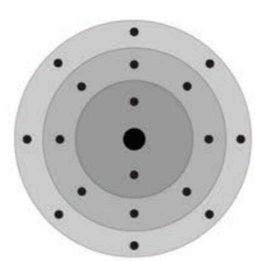


The atom you are looking for has

2

more Electrons than Oxygen (O).

The atom you are looking for is **directly above** of the atom with this Energy Level.



The atom you are looking for has

11

Protons in its Nucleus.

Atomic Number **11**
Sodium (Na)
Atomic Mass 22.99

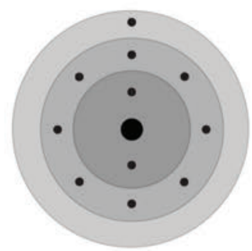
Na

The atom you are looking for has

11

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

12

Neutrons (usually) in its Nucleus.

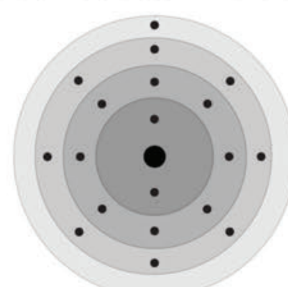
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
1 Electron on the **Third** Energy Level.

The atom you are looking for has

2

fewer Protons than Aluminum (Al).

The atom you are looking for is **directly above** the atom with this Energy Level.

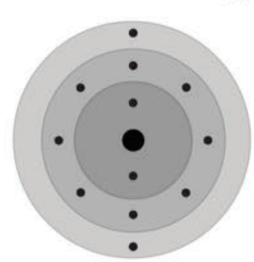


The atom you are looking for has

3

more Electrons than Oxygen (O).

The atom you are looking for is **directly to the left** of the atom with this Energy Level.



The atom you are looking for has

12

Protons in its Nucleus.

Atomic Number 12
Magnesium (Mg)
Atomic Mass 24.31

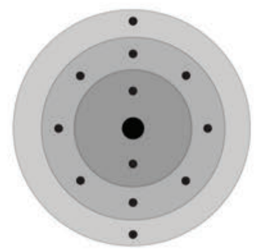
Mg

The atom you are looking for has

12

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

12

Neutrons (usually) in its Nucleus.

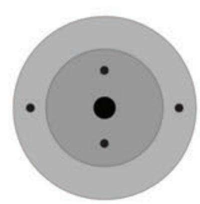
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
2 Electrons on the **Third** Energy Level.

The atom you are looking for has

10

more Protons than Helium (He).

The atom you are looking for is **directly below** the atom with this Energy Level.

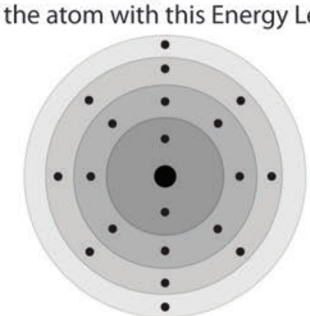


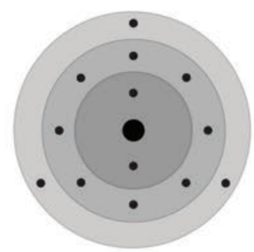
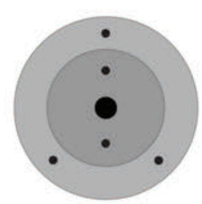
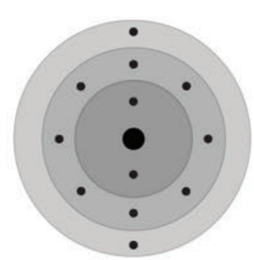
The atom you are looking for has

8

more Electrons than Beryllium (Be).

The atom you are looking for is **directly above** of the atom with this Energy Level.



The atom you are looking for has **13** Protons in its Nucleus.	Atomic Number **13** Aluminum (Al) Atomic Mass 26.98 **Al**
The atom you are looking for has **13** Electrons surrounding its Nucleus.	The atom you are looking for has this Energy Level Model: 
The atom you are looking for has **14** Neutrons (usually) in its Nucleus.	The atom you are looking for has **2** Electrons on the **First** Energy Level, **8** Electrons on the **Second** Energy Level, and **3** Electrons on the **Third** Energy Level.
The atom you are looking for has **8** more Protons than Boron (B).	The atom you are looking for is **directly below** the atom with this Energy Level. 
The atom you are looking for has **5** more Electrons than Oxygen (O).	The atom you are looking for is **directly to the right** of the atom with this Energy Level. 

The atom you are looking for has

14

Protons in its Nucleus.

Atomic Number **14**
Silicon (Si)
Atomic Mass 28.09

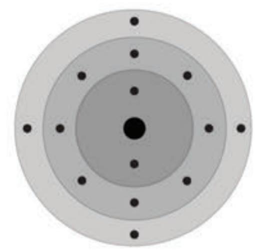
Si

The atom you are looking for has

14

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

14

Neutrons (usually) in its Nucleus.

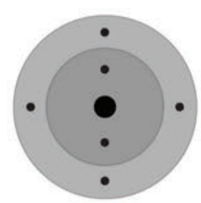
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
4 Electrons on the **Third** Energy Level.

The atom you are looking for has

3

fewer Protons than Chlorine (Cl).

The atom you are looking for is **directly below** the atom with this Energy Level.

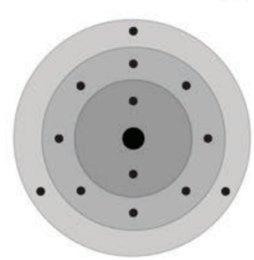


The atom you are looking for has

2

more Electrons than Magnesium (Mg).

The atom you are looking for is **directly to the right** of the atom with this Energy Level.



The atom you are looking for has

15

Protons in its Nucleus.

Atomic Number **15**
Phosphorous (P)
Atomic Mass 30.97

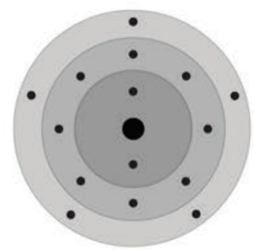
P

The atom you are looking for has

15

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

16

Neutrons (usually) in its Nucleus.

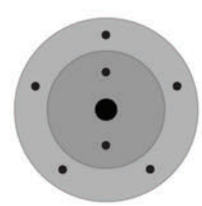
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
5 Electrons on the **Third** Energy Level.

The atom you are looking for has

8

more Protons than Nitrogen (N).

The atom you are looking for is **directly below** the atom with this Energy Level.

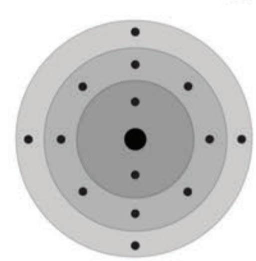


The atom you are looking for has

3

fewer Electrons than Argon (Ar).

The atom you are looking for is **directly to the right** of the atom with this Energy Level.



The atom you are looking for has

16

Protons in its Nucleus.

Atomic Number **16**
Sulfur (S)
Atomic Mass 32.07

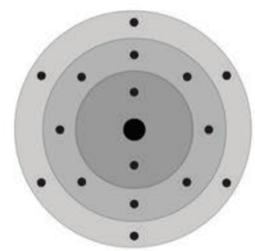
S

The atom you are looking for has

16

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

16

Neutrons (usually) in its Nucleus.

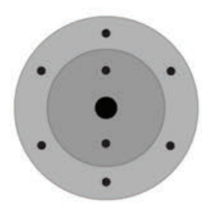
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
6 Electrons on the **Third** Energy Level.

The atom you are looking for has

10

more Protons than Carbon (C).

The atom you are looking for is **directly below** the atom with this Energy Level.

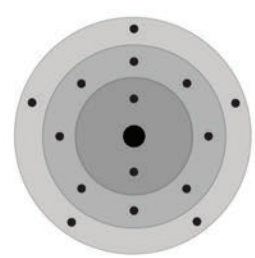


The atom you are looking for has

6

more Electrons than Neon (Ne).

The atom you are looking for is **directly to the right** of the atom with this Energy Level.



The atom you are looking for has

17

Protons in its Nucleus.

Atomic Number **17**
Chlorine (Cl)
Atomic Mass 35.45

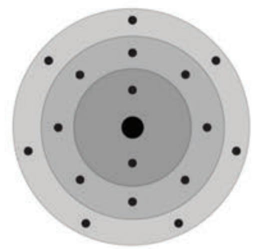
Cl

The atom you are looking for has

17

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

18

Neutrons (usually) in its Nucleus.

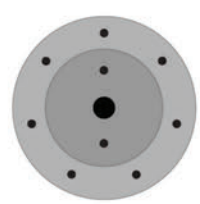
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
7 Electrons on the **Third** Energy Level.

The atom you are looking for has

3

fewer Protons than Calcium (Ca).

The atom you are looking for is **directly below** the atom with this Energy Level.

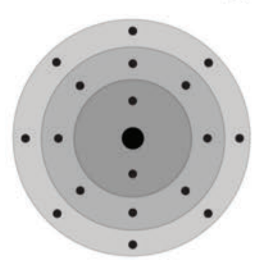


The atom you are looking for has

8

more Electrons than Fluorine (F).

The atom you are looking for is **directly to the left** of the atom with this Energy Level.



The atom you are looking for has

18

Protons in its Nucleus.

Atomic Number **18**
Argon (Ar)
Atomic Mass 39.95

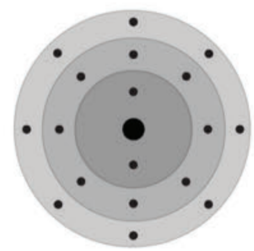
Ar

The atom you are looking for has

18

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

22

Neutrons (usually) in its Nucleus.

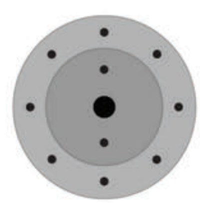
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
8 Electrons on the **Third** Energy Level.

The atom you are looking for has

7

more Protons than Sodium (Na).

The atom you are looking for is **directly below** the atom with this Energy Level.

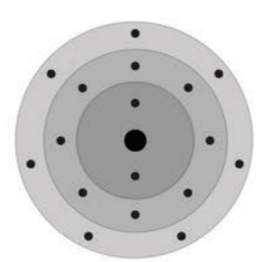


The atom you are looking for has

8

more Electrons than Neon (Ne).

The atom you are looking for is **directly to the right** of the atom with this Energy Level.



The atom you are looking for has

19

Protons in its Nucleus.

Atomic Number **19**
Potassium (K)
Atomic Mass 39.10

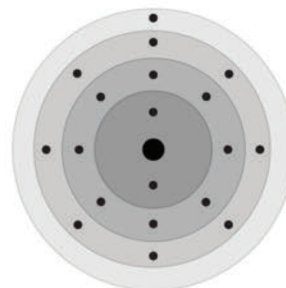
K

The atom you are looking for has

19

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

20

Neutrons (usually)
in its Nucleus.

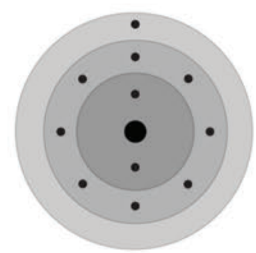
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level,
8 Electrons on the **Third** Energy Level, and
1 Electron on the **Fourth** Energy Level.

The atom you are looking for has

4

more Protons
than Phosphorous (P).

The atom you are looking for is
directly below the atom with this Energy Level.

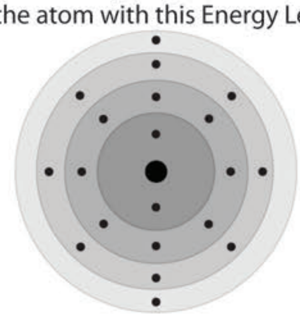


The atom you are looking for has

18

more Electrons
than Hydrogen (H).

The atom you are looking for is
directly to the left
of the atom with this Energy Level.



The atom you are looking for has

20

Protons in its Nucleus.

Atomic Number **20**
Calcium (Ca)
Atomic Mass 40.08

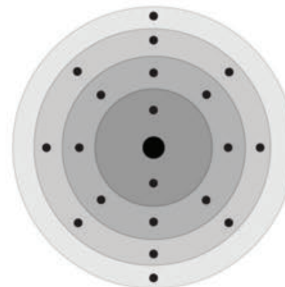
Ca

The atom you are looking for has

20

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

20

Neutrons (usually) in its Nucleus.

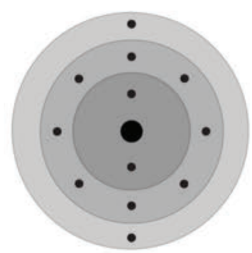
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level,
8 Electrons on the **Third** Energy Level, and
2 Electrons on the **Fourth** Energy Level.

The atom you are looking for has

8

more Protons than Magnesium (Mg).

The atom you are looking for is **directly below** the atom with this Energy Level.

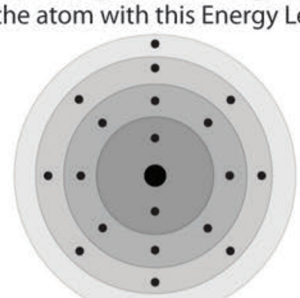


The atom you are looking for has

2

more Electrons than Argon (Ar).

The atom you are looking for is **directly to the right** of the atom with this Energy Level.



Chapter 4, Lesson 3: The Periodic Table and Energy-Level Models

Key Concepts

- The electrons surrounding an atom are located in regions around the nucleus called "energy levels".
- An energy level represents the 3-dimensional space surrounding the nucleus where electrons are most likely to be.
- The first energy level is closest to the nucleus. The second energy level is a little farther away than the first. The third is a little farther away than the second, and so on.
- Each energy level can accommodate or "hold" a different number of electrons before additional electrons begin to go into the next level.
- When the first energy level has 2 electrons, the next electrons go into the second energy level until the second level has 8 electrons.
- When the second energy level has 8 electrons, the next electrons go into the third energy level until the third level has 8 electrons.
- When the third energy level has 8 electrons, the next 2 electrons go into the fourth energy level.
- The electrons in the energy level farthest from the nucleus are called *valence* electrons.
- Atoms in the same column (group) in the periodic table have the same number of valence electrons.

Summary

Students will again focus on the first 20 elements. Students will first look at a diagram and animation to understand the basic pattern of the arrangement of electrons on energy levels around an atom. Students will be given cards with information about the electrons and energy levels for each of the first 20 atoms. They will again try to correctly match the cards with each element.

Objective

Students will be able to interpret the information given in the periodic table to describe the arrangement of electrons on the energy levels around an atom.

Evaluation

The activity sheet will serve as the "Evaluate" component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding.

About this Lesson

Be sure that the 20 atom name cards are posted around the room. You will need the five cards on the right hand side of each sheet. This lesson is intended as a follow-up to chapter 4, lesson 2.

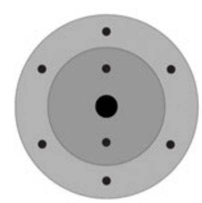
ENGAGE

1. Introduce students to the idea that electrons surround the nucleus of an atom in regions called energy levels.

Review with students that in lesson two they focused on the number of protons, neutrons, and electrons in the atoms in each element. In this lesson, they will focus on the arrangement of the electrons in each element.

Project the image *Energy level cross-section*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson3#energy_level_cross_section
Explain to students that electrons surround the nucleus of an atom in three dimensions, making atoms spherical. They can think of electrons as being in the different energy levels like concentric spheres around the nucleus. Since it is very difficult to show these spheres, the energy levels are typically shown in 2 dimensions.

Project the image *Oxygen atom*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson3#oxygen_atom
Tell students that this energy level model represents an atom. The nucleus is represented by a dot in the center, which contains both protons and neutrons. The smaller dots surrounding the nucleus represent electrons in the energy levels. Let students know that they will learn more about electrons and energy levels later in this lesson.



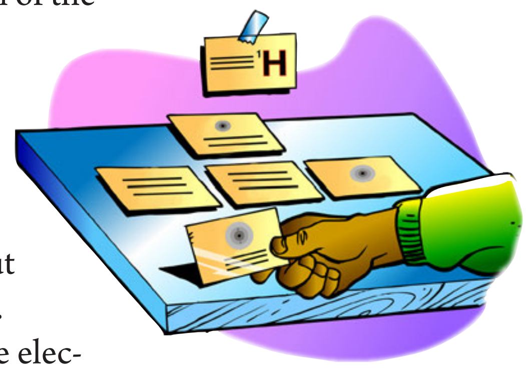
Have students look at the *Periodic table of the elements 1–20* they used in lesson 2 to answer the following question:

- **Can you identify which atom this model represents?**
  If students can't answer this question, point out that there are 8 electrons. Because neutral atoms in the periodic table have the same number of electrons as protons, the atom must have 8 protons. The number of protons is the same as the atomic number, so the atom is oxygen.

Read more about energy level models in the additional teacher background section at the end of this lesson.

2. **Have groups work together to place each card with its correct atom.**

 Show students that you have 80 cards (4 for each of the first 20 elements). Before distributing the cards, explain that each card contains information about electrons and energy levels for the first 20 elements of the periodic table. The students' job is to read the card carefully, figure out which element the card is describing, and put the card at the spot in the room for that element. Remind students that they will need to count the electrons in order to identify each atom. Once students understand what their assignment is, distribute the cards to groups.



3. **Discuss the placement of the cards for two or three atoms.**

 After all cards have been placed at the 20 different atoms, select two or three atoms and review whether the cards were placed correctly. This review will help reinforce the concepts about the structure of atoms and help students determine the number of protons and electrons in each atom.

 Give each student a *Periodic Table of Energy Levels* activity sheet. This table contains energy level models for the first 20 elements. The electrons are included only for the atoms at the beginning and end of each period.

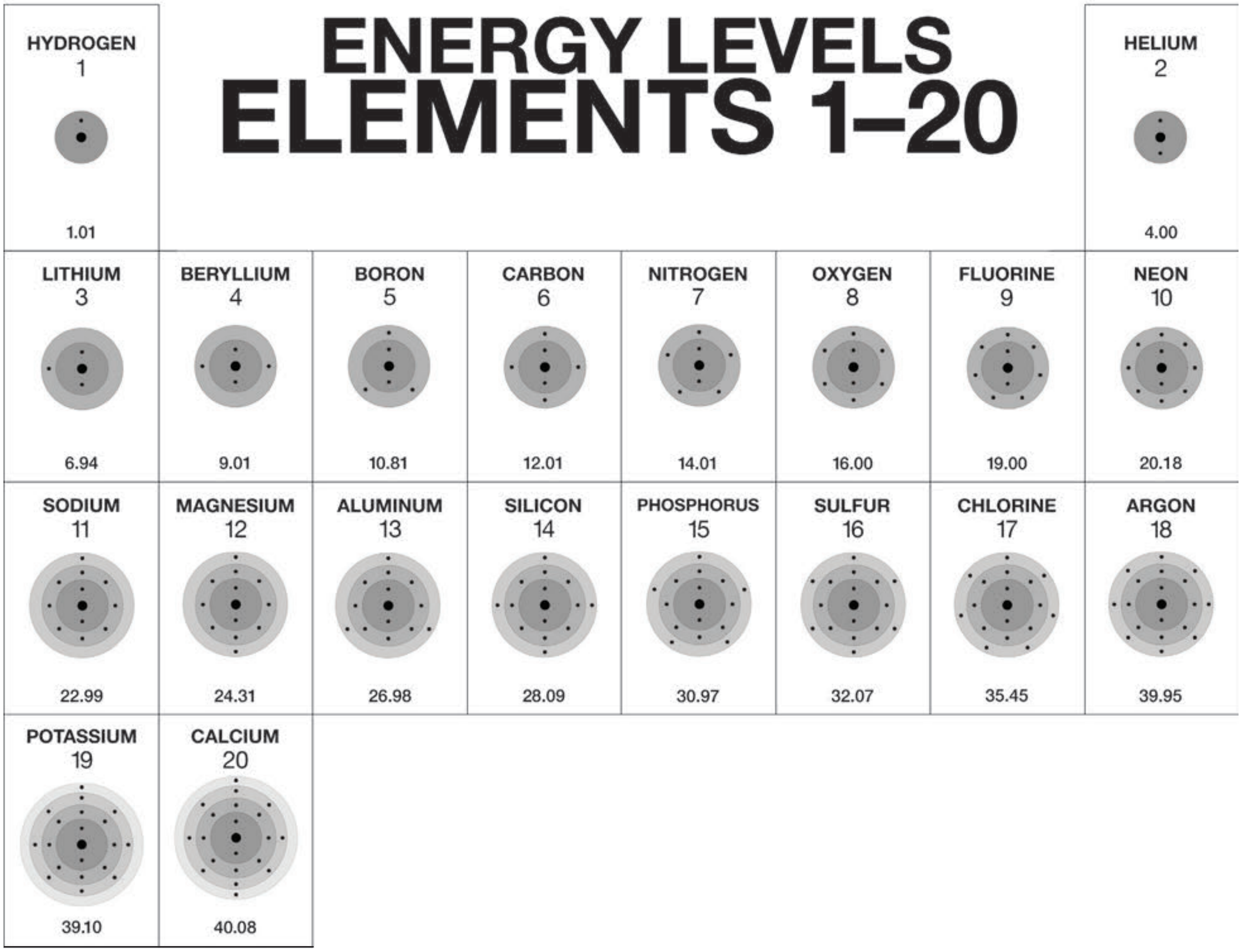
EXPLORE

4. **Project the *Periodic table of energy levels* and discuss the arrangement of electrons as students complete their activity sheet.**

 Project the image *Periodic table of energy levels*.
 www.middleschoolchemistry.com/multimedia/chapter4/lesson3#energy_levels

 The image you project contains all of the electrons for elements 1–20. However, the periodic table on the activity sheet contains electrons only for the elements at the beginning and end of each period. Discuss the arrangement of electrons within the energy levels for these atoms and have students fill in the electrons for the other atoms.

 Note: *In the energy level diagrams, the electrons are spread out evenly in the level. Some books show them spread out this way and some show them in pairs. The pairing of electrons is meant to represent that electrons are in separate orbitals within each energy level. At the middle school* *level, it is not necessary for students to learn about electron orbitals. This information is offered so that it is clearer to you why electrons are often shown in pairs in energy level diagrams and in the dot diagrams used as an extension at the end of this chapter. An orbital defines a region within an energy level where there is a high probability of finding a pair of electrons. There can be a maximum of two electrons in each orbital. This is why the electrons are often shown in pairs within an energy level.*



Tell students that the rows across on the periodic table are called *periods*.

Period 1

- *Hydrogen*
  Explain that hydrogen has 1 proton and 1 electron. The 1 electron is on the first energy level.
- *Helium*
  Explain that helium has 2 protons and 2 electrons. The 2 electrons are on the first energy level.

Period 2

- *Lithium*
  Explain that lithium has 3 protons and 3 electrons. There are 2 electrons on the first energy level and 1 electron on the second. Explain that the first energy level can only have 2 electrons so the next electron in lithium is on the next (second) level.
- *Neon*
  Explain that neon has 10 protons and 10 electrons. There are 2 electrons on the first energy level and 8 electrons on the second level.
- *Beryllium–fluorine*
  Help students fill in the correct number of electrons in the energy levels for the rest of the atoms in period 2.

Period 3

- *Sodium*
  Explain that sodium has 11 protons and 11 electrons. There are 2 electrons on the first energy level, 8 electrons on the second level, and 1 electron on the third energy level. Explain that the second energy level can only have 8 electrons so the next electron in sodium has to be on the next (third) level.
- *Argon*
  Explain that argon has 18 protons and 18 electrons. There are 2 electrons on the first energy level, 8 electrons on the second level, and 8 electrons on the third energy level. Have students complete the energy level model for argon in their periodic table.
- *Magnesium–chlorine*
  Help students fill in the correct number of electrons in the energy levels for the rest of the atoms in period 3.

Period 4

- *Potassium*
  Explain that potassium has 19 protons and 19 electrons. There are 2 electrons on the first energy level, 8 electrons on the second level, 8 electrons on the third energy level, and 1 on the fourth energy level. Explain that after the third energy level has 8 electrons, the next electron goes into the fourth level.
- *Calcium*
  Help students fill in the correct number of electrons in the energy levels for calcium.

***Note**: Students may wonder why an energy level can hold only a certain number of electrons. The answer to this is far beyond the scope of a middle school chemistry unit. It involves thinking of electrons as 3-dimensional waves and how they would interact with each other and the nucleus.*

5. Have students look for patterns in rows and columns of the first 20 elements in the periodic table.

Continue to project the image *Periodic table of energy levels for elements 1–20* and have students look at their activity sheets to find patterns in the number of electrons within each energy level.

Have students look at the periods (rows going across).
Number of energy levels in each period

- The atoms in the first period have electrons in 1 energy level.
- The atoms in the second period have electrons in 2 energy levels.
- The atoms in the third period have electrons in 3 energy levels.
- The atoms in the fourth period have electrons in 4 energy levels.

How the electrons fill in the energy levels

- First energy level = 1, 2
- Second energy level = 1, 2, 3, …8
- Third energy level = 1, 2, 3, …8
- Fourth energy level = 1, 2

Read more about the periodic table in the additional teacher background section at the end of this lesson.

A certain number of electrons go into a level before the next level can have electrons in it. After the first energy level contains 2 electrons (helium), the next electrons go into the second energy level. After the second energy level has 8 electrons (neon), the next electrons go into the third energy level. After the third energy level has 8 electrons (argon), the next 2 electrons go into the fourth energy level.

***Note**: The third energy level can actually hold up to 18 electrons, so it is not really filled when it has 8 electrons in it. But when the third level contains 8 electrons, the next 2 electrons go into the fourth level. Then, believe it or not, 10 more electrons continue to fill up the rest of the third level. Students do not need to know this.*

Have students look at the groups (columns going down).
Tell students that the vertical columns in the periodic table are called *groups* or *families*. Ask students to compare the number of electrons in the outermost energy level for the atoms in a group. Students should realize that each atom in a group has the same number of electrons in its outermost energy level. For instance, hydrogen, lithium, sodium, and potassium all have 1 electron on their outer energy level. Let students know that these electrons in the outermost energy level are called *valence* electrons. They are the electrons responsible for bonding, which students will investigate in the next lesson.

EXTEND

6. Compare the way different elements react chemically and relate this to their location on the periodic table.

Tell students that in the periodic table atoms in the same column, called a group, share certain characteristics and can react in a similar way.

Project the video *Sodium in water and potassium in water*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson3#sodium_in_water
www.middleschoolchemistry.com/multimedia/chapter4/lesson3#potassium_in_water
Students will see that although potassium reacts more vigorously than sodium, the reactions are similar. Have students look at the periodic table to see where sodium and potassium are in relation to one another.

Project the video *Calcium in water*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson3#calcium_in_water
Students will see that this reaction is different from the sodium and the potassium. Have them locate calcium on the periodic table and point out that it is in a different group than sodium and potassium.

Project the videos *Sodium in acid and potassium in acid*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson3#sodium_in_acid
www.middleschoolchemistry.com/multimedia/chapter4/lesson3#potassium_in_acid
Show sodium reacting with acid and then potassium reacting with acid. The HCl is hydrochloric acid. The HNO_3 is nitric acid. Each acid is used in two different concentrations. Make sure students realize that the sodium and potassium react in a similar way even though the potassium reacts more vigorously.

Project the video *Calcium in acid*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson3#calcium_in_acid
Point out that calcium reacts differently from the sodium and the potassium.

Ask students:

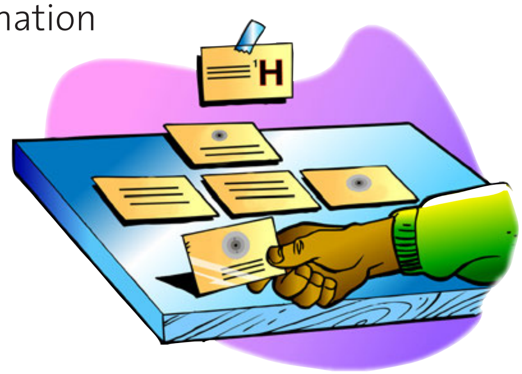
- **Do elements in the same group have similar properties and react in similar ways?**
  Students should realize that sodium and potassium are in the same group and react similarly. Calcium is near them on the periodic table, but is in a different group, so it reacts differently.

Activity Sheet
Chapter 4, Lesson 3
The Periodic Table and Energy Level Models

Name ____________________

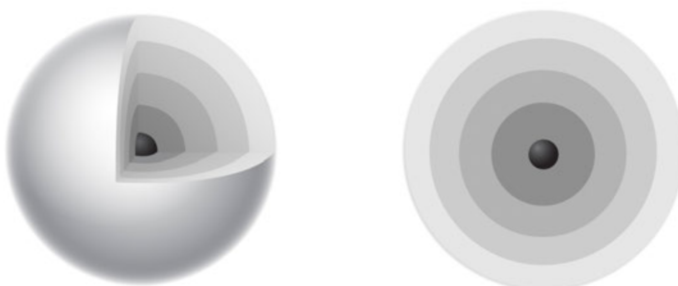
Date ____________________

Your group will receive a set of cards with information about the energy levels of a particular atom. Your job is to figure out which atom the card describes and to place it in the area in your classroom for that atom. Use the activity sheet from lesson 2 along with this activity sheet as a reference.



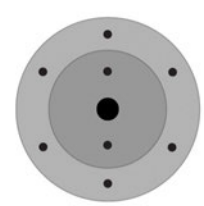
Energy levels

Electrons surround the nucleus of an atom in regions called *energy levels*. Even though atoms are spherical, the energy levels in an atom are more easily shown in concentric circles.



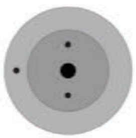
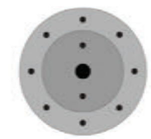
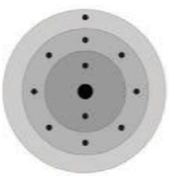
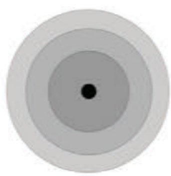
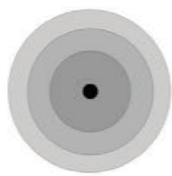
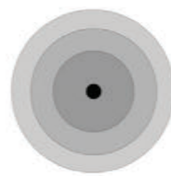
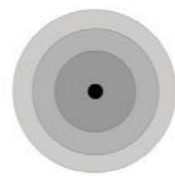
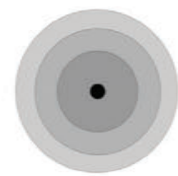
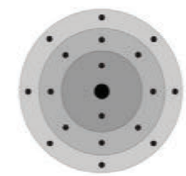
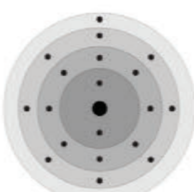
Which atom is this supposed to be?

The larger dot in the center of this atom represents the nucleus, which contains both protons and neutrons. The smaller dots surrounding the nucleus represent electrons. In order to figure out which atom this represents, count up the number of electrons. There are 8 electrons in this atom. Because the number of electrons and protons is the same in an atom, this atom has 8 protons. Look at the chart Periodic Table, Elements 1–20. The number of protons is the same as the atomic number, so this drawing represents an oxygen atom.



ENERGY LEVELS ELEMENTS 1–20

Complete each energy level model by drawing the correct number of electrons in their corresponding energy levels.

HYDROGEN 1  1.01							**HELIUM** 2  4.00
LITHIUM 3  6.94	**BERYLLIUM** 4  9.01	**BORON** 5  10.81	**CARBON** 6  12.01	**NITROGEN** 7  14.01	**OXYGEN** 8  16.00	**FLUORINE** 9  19.00	**NEON** 10  20.18
SODIUM 11  22.99	**MAGNESIUM** 12  24.31	**ALUMINUM** 13  26.98	**SILICON** 14  28.09	**PHOSPHORUS** 15  30.97	**SULFUR** 16  32.07	**CHLORINE** 17  35.45	**ARGON** 18  39.95
POTASSIUM 19  39.10	**CALCIUM** 20  40.08						

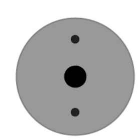
Additional Teacher Background
Chapter 4 Lesson 3, p. 291

As the note on page 292 points out, there are other ways to model the electron energy levels of atoms. Some middle school texts show the electrons in pairs on an energy level. This pairing of electrons is intended to suggest information about the substructure *within* energy levels. This substructure is made up of regions called *orbitals* which comprise each energy level. The shape and size of the orbital is defined by the space around the nucleus where there is a high probability of finding electrons. There can be a maximum of two electrons in any orbital so showing electrons in pairs on an energy level model is an attempt to suggest information about the orbitals within the level.

In Middle School Chemistry, we chose to spread electrons out evenly on energy levels to indicate only the *number* of electrons on a level and not to suggest anything about the substructure of orbitals *within* energy levels. An understanding that the different energy levels can accommodate a certain number of electrons seems enough for students in middle school. They will see more refined models in high school and college when they learn more details about the orbitals within energy levels.

Some teachers might like to use a different model that shows more details of orbitals because it is more complete, even if they do not intend to explain those aspects of the model in much detail. Another argument is that a model showing paired and unpaired electrons may be useful for certain discussions about bonding. Other teachers may be more comfortable showing a less-detailed model even if it leaves out certain aspects of energy levels because they do not intend to discuss those details and they intend to handle bonding in a more general way. No model can be complete and accurate for all purposes and all have limitations. All models involve aspects of judgment and compromise. A good model focuses on the important points without too much to distract from those main features. The model you choose will have a lot to do with how much you think is important to explain and what the students are able to understand.

Some energy level models you might see and what they represent
For helium (atomic number 2), the energy level model in Middle School Chemistry is:

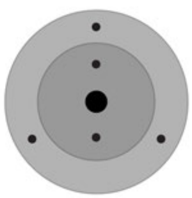


Helium has two electrons on the first energy level.

Some other middle school texts might show an energy level model for helium like this:

The *first* energy level has only one orbital. This is known as the 1s orbital. The "1" means that it is in the first energy level and the "s" stands for an orbital within that energy level with a particular shape. This 1s orbital can hold up to two electrons. So helium has its two electrons in the 1s orbital. The practice of showing the electrons together or *paired* in an energy level is meant to indicate how many orbitals in that level have been completely occupied by two electrons. For the first energy level, the pairing is not very useful for showing which orbitals are full and which aren't because there is only one orbital. But it becomes more useful for atoms that have more orbitals where some orbitals may be filled and others not.

For boron (atomic number 5), the energy level model in Middle School Chemistry is:



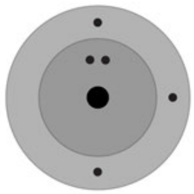
Boron has 2 electrons on the first energy level and 3 electrons on the second level.

Some other middle school texts might show an energy level model for boron like this:

The model shows that boron has two electrons in the 1s orbital of the first energy level which are shown as paired. It also has 3 electrons in the second energy level.

The second energy level is made up of four orbitals. There is a spherical orbital called 2s. The "2" means that it is in the second energy level. It is like the 1s orbital but is further from the nucleus. The second energy level also has 3 other orbitals that are all the same shape and distance from the nucleus but oriented in different directions. These orbitals are called 2p. The "p" orbitals are a different shape than the "s" orbitals. The 2s orbital can hold up to two electrons and each of the 2p orbitals can also hold up to 2 electrons. So the second energy level can hold up to eight electrons in its four orbitals. In this model of boron, two electrons are shown as paired in the 2s orbital and the last electron is shown in one of the 2p orbitals.

Another middle school text might show a model of boron like this:



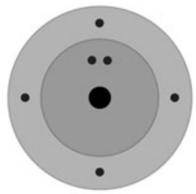
Here, they paired the electrons in the 1s orbital but did not show the detail of pairing the electrons in the 2s orbital of the second energy level. They chose to spread the three electrons out on the second energy level.

For carbon (atomic number 6), the energy level model in Middle School Chemistry is:

Carbon has 2 electrons on the first energy level and 4 on the second. Some other middle school texts might show a model of carbon like this:

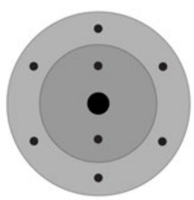
This model shows that carbon has two electrons in the 1s orbital of the first energy level which are shown as paired. It also has 4 electrons in the second energy level. In this model, two electrons are shown as paired in the 2s orbital and the other two electrons are shown separately or unpaired. This is done to indicate that each of the electrons is in a separate 2p orbital. One of the details of orbitals is that an electron goes into an empty available orbital of the same type before it goes into an orbital that already has an electron in it.

Another middle school text might show a model of carbon like this:



This model pairs the 1s electrons but spreads out the four electrons in the second energy level regardless of what orbital they are in. This approach would show electrons being paired on the second energy level for the first time in nitrogen.

For oxygen (atomic number 8), the energy level model in Middle School Chemistry is:



Oxygen has 2 electrons on the first energy level and 6 on the second. Oxygen is an interesting example because the other two types of models come out with the same result which looks like this:

Here, the electrons are paired in the 1s orbital. In the second energy level, whether the electrons are paired in the 2s to begin with or whether they are spread out and only paired after placing 1 electron in each of the four orbitals and then adding the last two electrons to make two pairs, the result is the same.

If the energy level models in Middle School Chemistry are different than those in your text book, you can use either one to teach that energy levels only have a certain number of electrons. You could also use the difference to suggest that there is more detail about energy levels that students may learn about later.

Additional Teacher Background
Chapter 4 Lesson 3, p. 295

What determines the shape of the standard periodic table?
One common question about the periodic table is why it has its distinctive shape. There are actually many different ways to represent the periodic table including circular, spiral, and 3-D. But in most cases, it is shown as a basically horizontal chart with the elements making up a certain number of rows and columns. In this view, the table is not a symmetrical rectangular chart but seems to have steps or pieces missing.

The key to understanding the shape of the periodic table is to recognize that the characteristics of the atoms themselves and their relationships to one another determine the shape and patterns of the table.

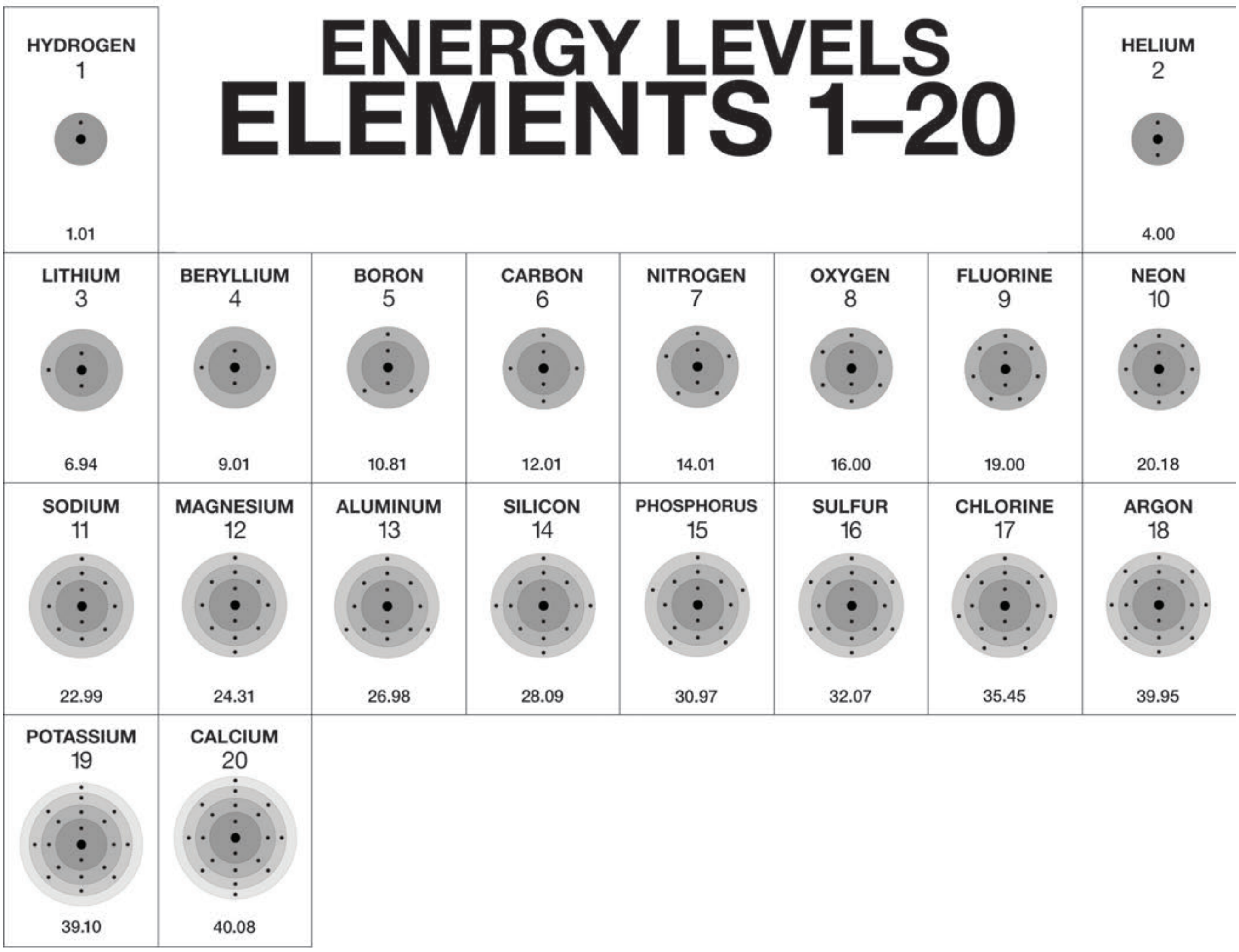
The Periodic Table of the Elements

3 Li Lithium 6.94	
3	Atomic Number
Li	Element Symbol
Lithium	Element Name
6.94	Average Atomic Mass

1 H Hydrogen 1.01																	2 He Helium 4.00
3 Li Lithium 6.94	4 Be Beryllium 9.01											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.39	31 Ga Gallium 69.72	32 Ge Germanium 72.61	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybednum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.91	56 Ba Barium 137.33	57 La Lanthanum 138.91	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.84	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.98	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium 178.49	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (269)	109 Mt Meitnerium (268)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (272)	112 Cn Copernicium (285)						

58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.97
90 Th Thorium 232.04	91 Pa Proactinium 231.04	92 U Uranium 238.03	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einstenium (252)	100 Fm Fermium (257)	101 Md Mendelevium 168.93	102 No Nobelium (259)	103 Lr Lawrencium (262)

A helpful starting point for explaining the shape of the periodic table is to look closely at the structure of the atoms themselves. You can see some important characteristics of atoms by looking at the chart of energy level diagrams. Remember that an energy level is a region around an atom's nucleus that can hold a certain number of electrons. The chart shows the number of energy levels for each element as concentric shaded rings. It also shows the number of protons (atomic number) for each element under the element's name. The electrons, which equal the number of protons, are shown as dots within the energy levels. The relationship between atomic number, energy levels, and the way electrons fill these levels determines the shape of the standard periodic table.



What determines the sequence of the elements?

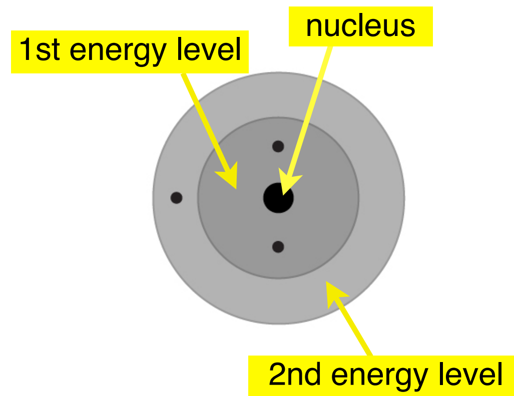
One of the main organizing principles of the periodic table is based on the atomic number (number of protons in the nucleus) of the atoms. If you look at any row, the atoms are arranged in sequence with the atomic number increasing by one from left to right. Since the number of electrons equals the number of protons, the number of electrons also increases by one from left to right across a row.

What do the rows represent?

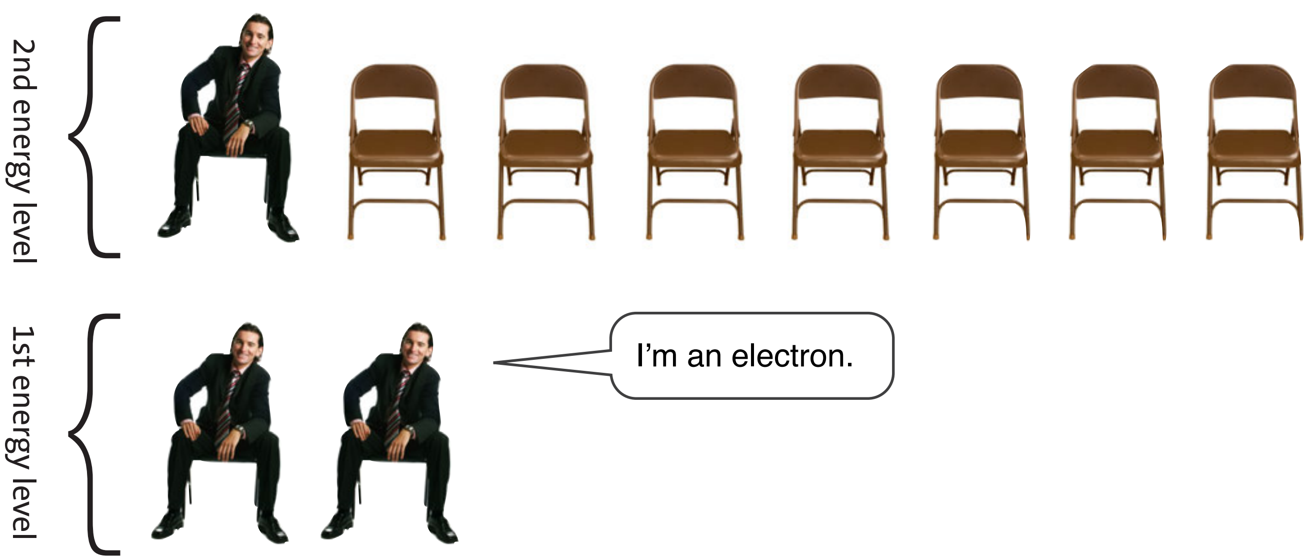
The rows in the periodic table correspond to the number of energy levels of the atoms in that row. If you look at the chart, you can see that the atoms in the first row have one energy level. The atoms in the second row have two energy levels and so on. Understanding how electrons are arranged within the energy levels can help explain why the periodic table has as many rows and columns as it does. Let's take a closer look.

Electrons and Energy Levels

Every atom contains different energy levels that can hold a specific number of electrons. For a moment, let's imagine the simplest possible scenario: once all the positions are occupied within one energy level, any remaining electrons begin filling positions in the next energy level.



To picture this, imagine people filling rows of chairs in an auditorium. If each person sits next to another person until one row is filled, any remaining people must begin taking their seats in the second row, and so on.

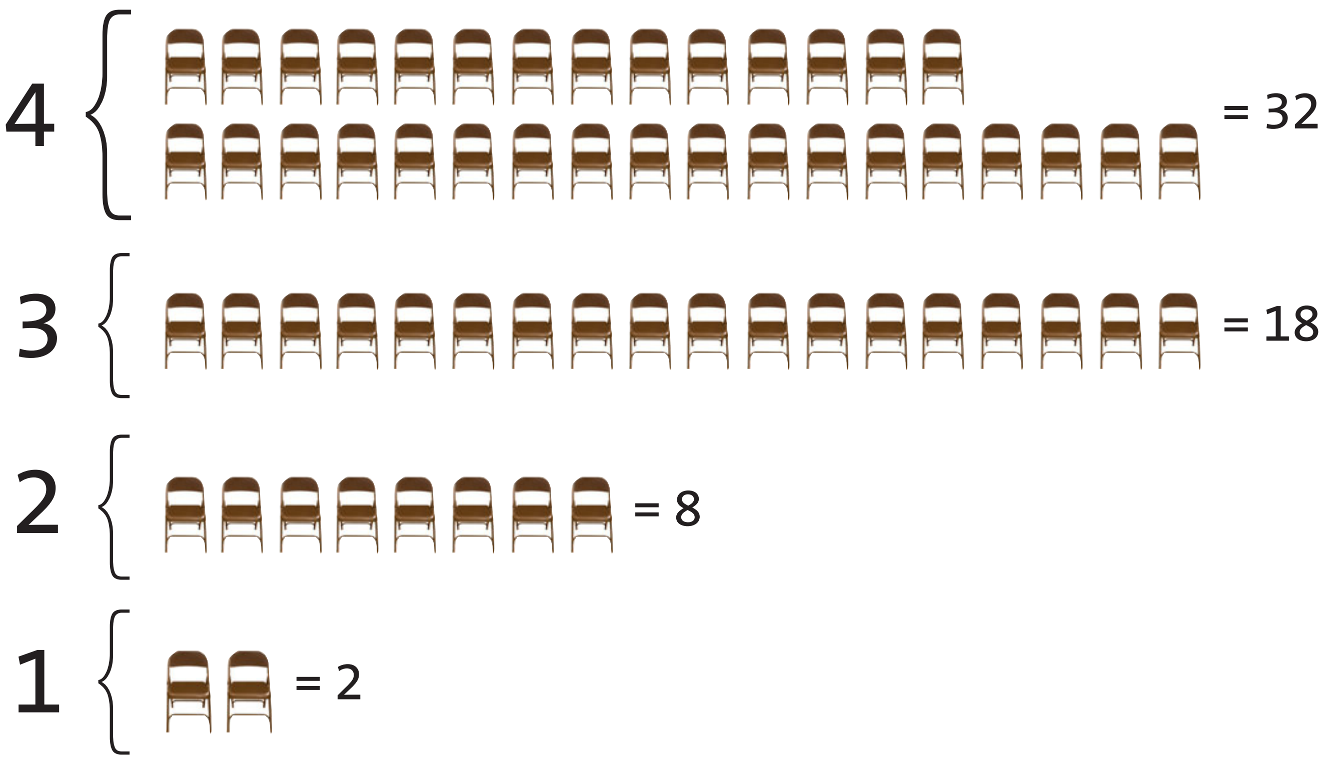


Not so bad, right? In general, this simple case is a helpful analogy. Electrons fill a given section until it is full, and then any more electrons move on to another unoccupied section where they continue filling there. Electrons begin filling the lowest energy level (closest to the nucleus) and then move on to higher energy levels (further form the nucleus). Unfortunately, the actual process is a bit more complicated. Let's see why.

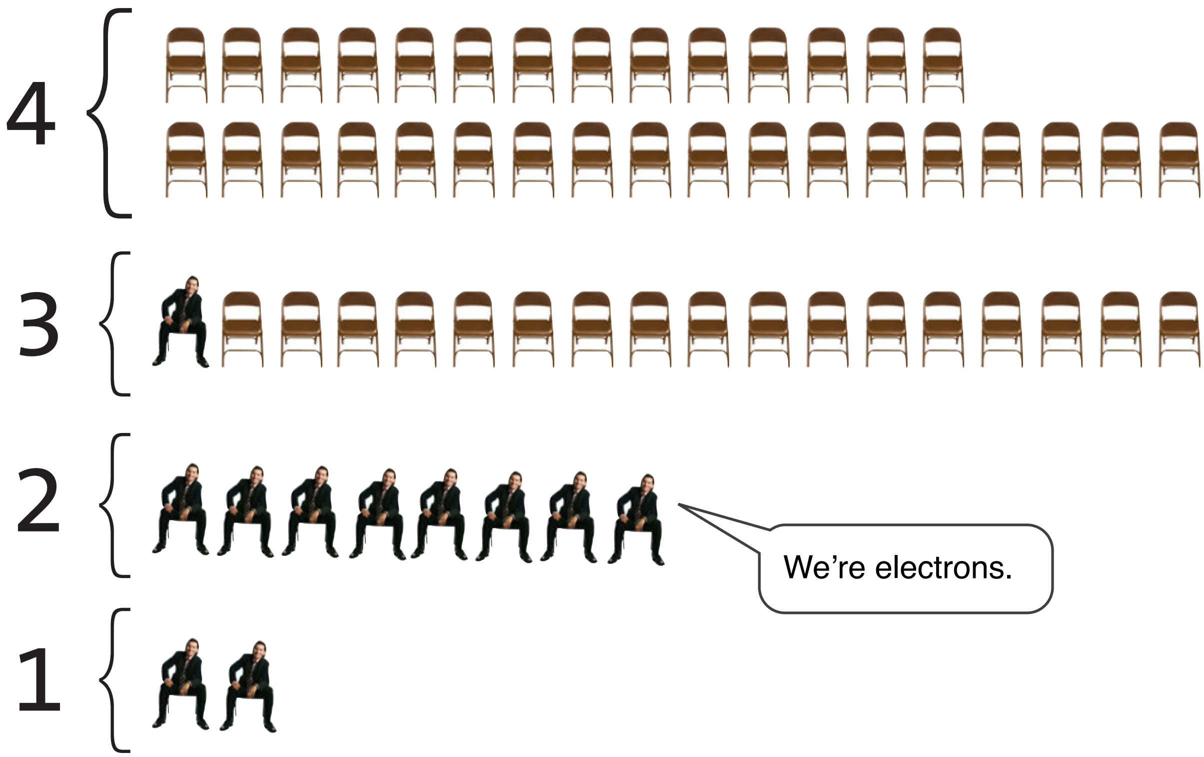
Energy Levels Can Hold Different Numbers of Electrons

One thing that is slightly tricky about electrons filling these energy levels is that not all the energy levels can hold the same number of electrons. While the first energy level can hold only 2 electrons, the second energy level can hold 8, the third can hold 18, and the fourth can hold 32. We'll stop there for now.

If we return to our rows of chairs analogy, it would be as if the first row was shorter than the second or third or fourth rows, so that after 2 people, any people remaining would have to begin occupying the second row. Then, if the second row were longer than the first row (but shorter than the third row), after 8 more people had been seated, any remaining individuals would have to begin occupying the third row.



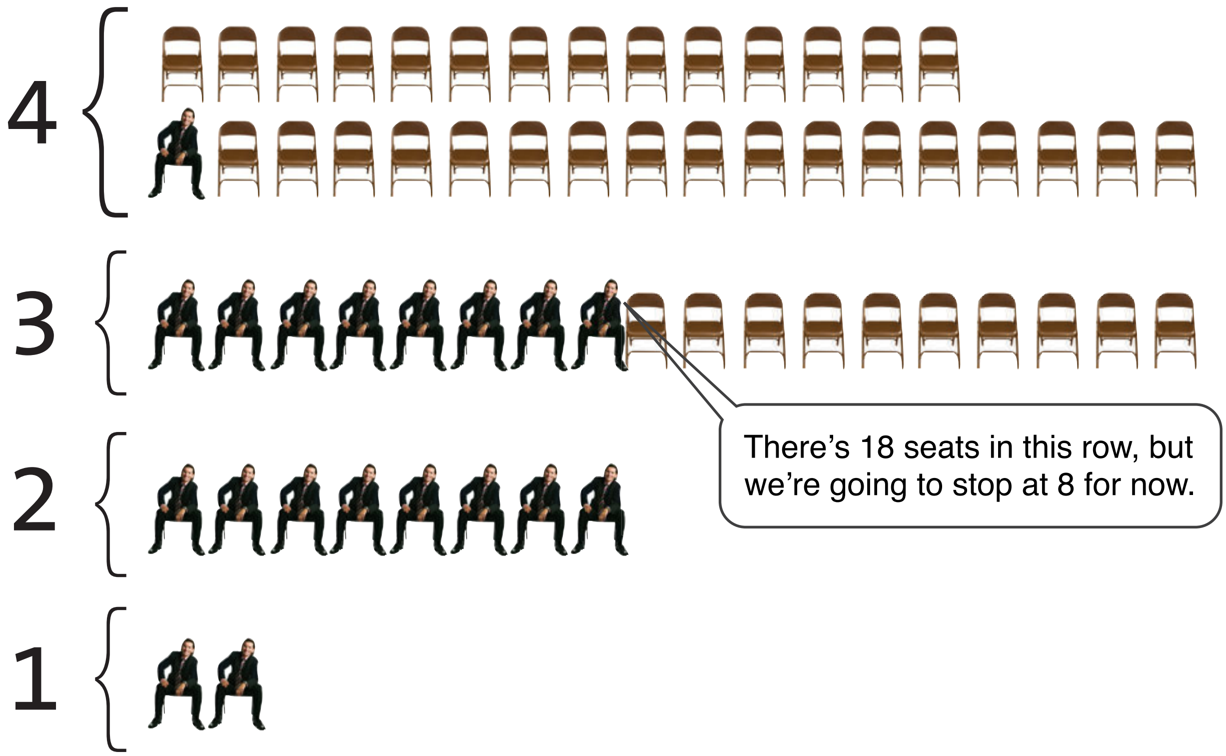
Extending our analogy of theater patrons as electrons, let's look at how the element sodium, with its 11 electrons, might fill these energy levels.



Because sodium has 11 electrons, it fills up the first energy level, which can hold only 2 electrons. It also fills up the second energy level, because it can only hold 8. Together, the first and second energy levels can hold a total of 10 electrons. Sodium has 11 electrons, so that final remaining electron that can't be accommodated by the first and second energy level begins filling in the third energy level. This pattern generally holds for the first 18 elements, up through argon, which has 18 electrons.

Energy Levels are Further Divided into Sections

But something funny happens beginning with potassium. Potassium has 19 electrons. Because the first, second, and third energy levels can hold a total of 28 electrons (2+8+18=28) it would seem that all the electrons of potassium could be "seated" within the third energy level. It turns out, however, that even though the third energy level has a total capacity of 18, only 8 "seats" are filled before the electrons begin filling the fourth energy level. So, potassium would fill up the energy levels like this:

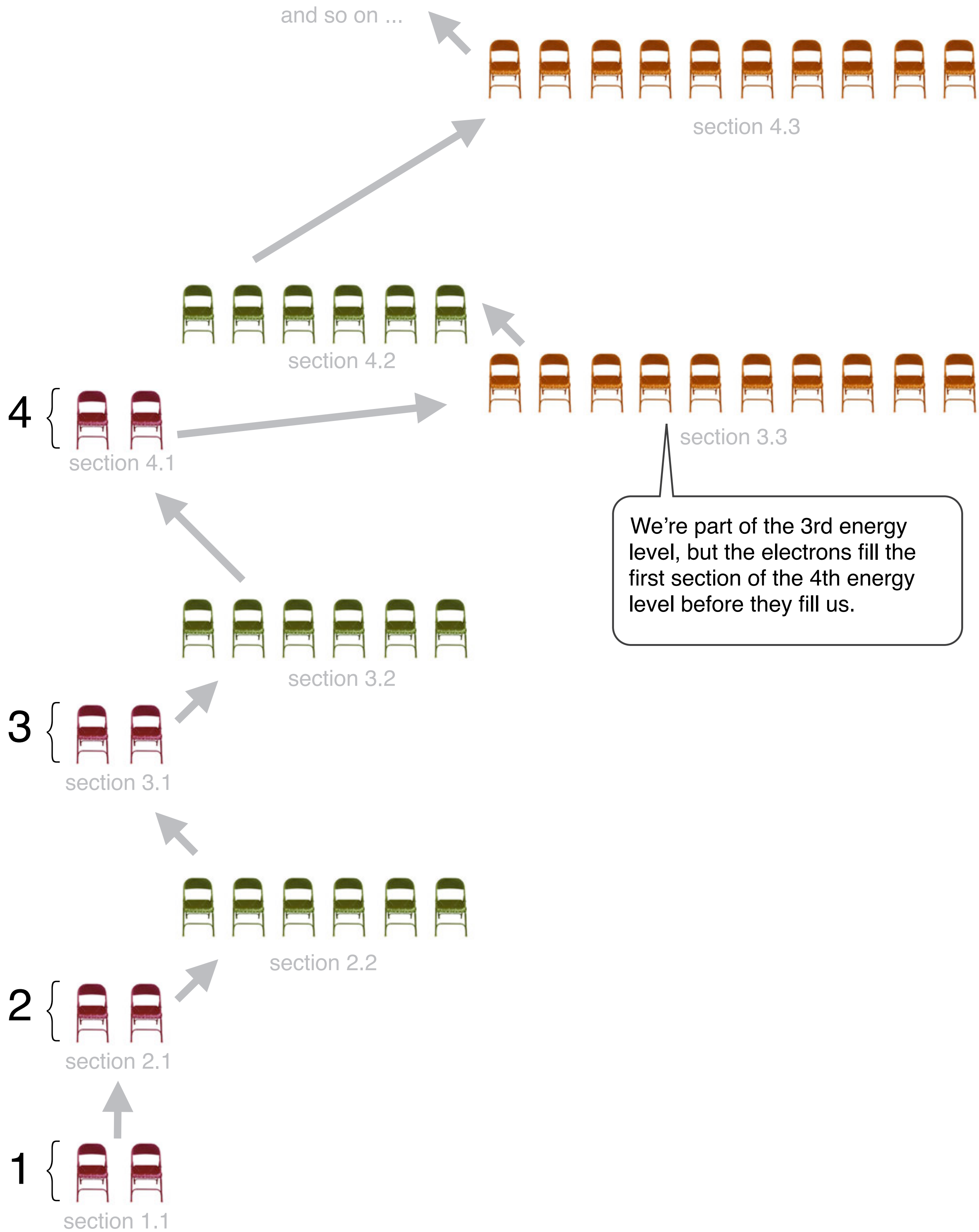


Whoa. *Whoa*. That's crazy. Why does *that* happen?

This is the second complication with our simple chairs analogy. It turns out that in addition to distinct energy levels (first, second, third, etc.) each energy level is further divided into sections where electrons can be found.

In terms of our analogy, the first row would have just one section. The second row would have two sections. The third row would have 3 sections and the fourth row would have four sections. As you can see, the number of sections an energy level has is equal to the number of that energy level.

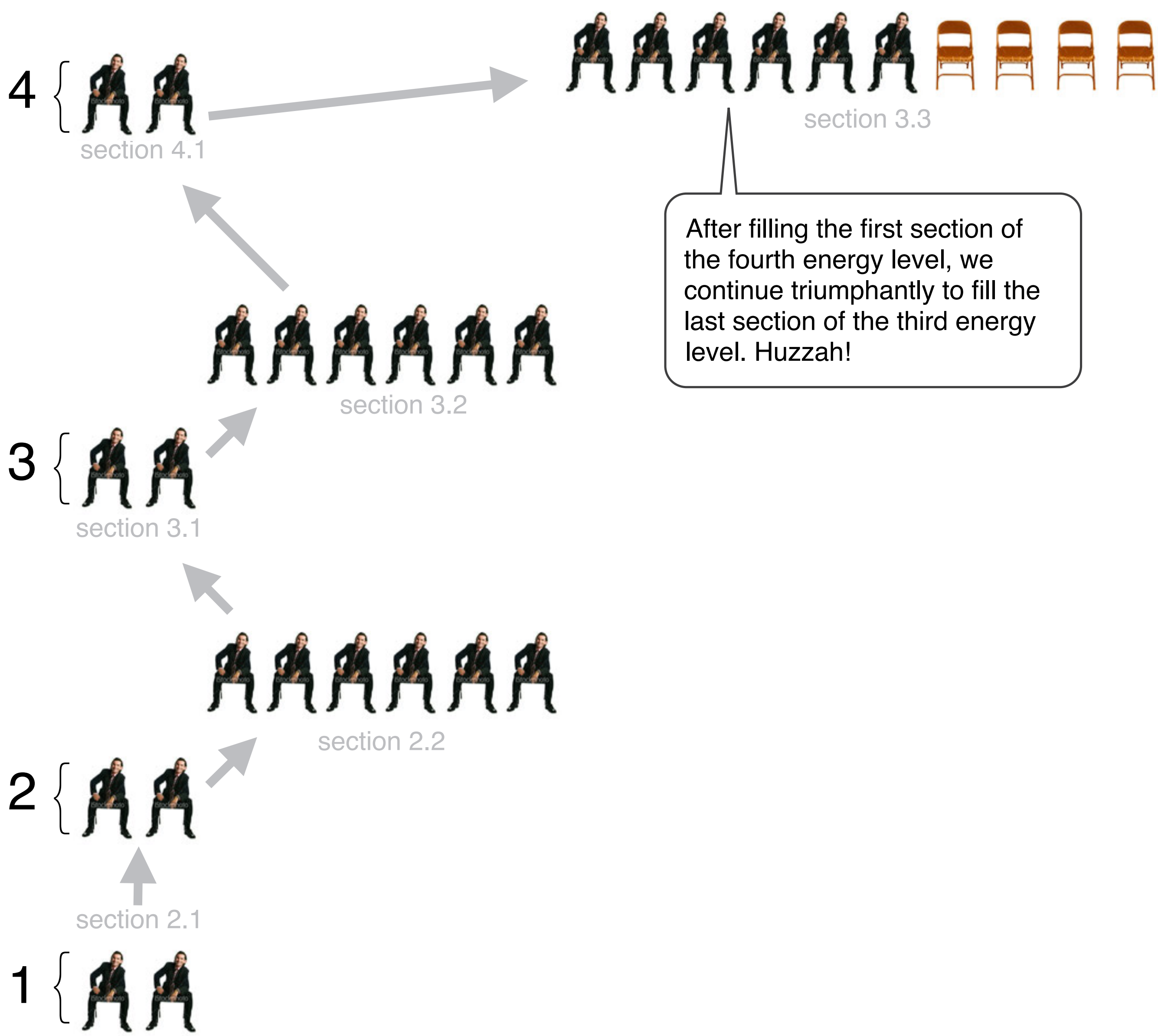
The reason why that last electron from potassium begins filling the fourth energy level rather than continuing to fill the third energy level is that the first section of the fourth energy level is actually closer (or at lower energy) than the last section of the third energy level (the last 10 "seats"). So, really, our chairs would now look something like this:


and so on ...
section 4.3
section 4.2
4
section 4.1
section 3.3
We're part of the 3rd energy level, but the electrons fill the first section of the 4th energy level before they fill us.
section 3.2
3
section 3.1
section 2.2
2
section 2.1
1
section 1.1

Admittedly, this doesn't look much like rows of charis in an auditoruim anymore, but the idea is still the same. Electrons will continue filling energy levels, one section at at time, until all the electrons are used up. When one section of the next energy level is actually lower in energy than the next section of the same energy level, the electrons will begin filling there. This is what we depicted in the diagram for potassium. Its last electron filled the first section in the fourth energy level, because that section was actually closer (at lower energy) than the last section of the third energy level.

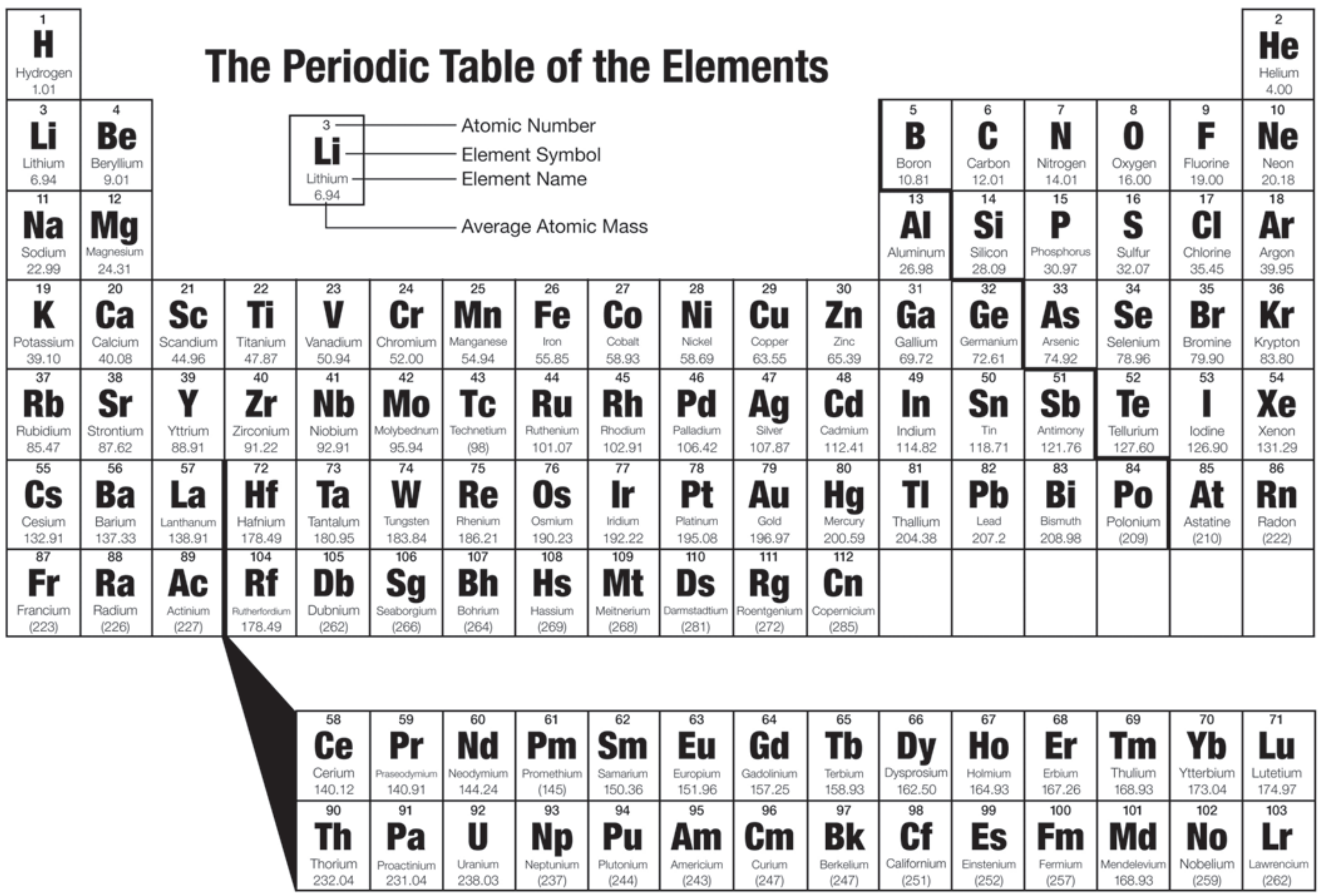
Eventually, the electrons will continue filling the empty section in the third energy level. The idea is exactly what we've just described. Unusual as it might seem, in some cases, the first section of the next energy level is filled before the electrons continue to fill the last section of the preceding energy level.

Consider, for example, the element Iron. Its 26 electrons would fill energy levels like this:



Whew! So what does all of this mean?

Mainly this: understanding how electrons fill energy levels can help us to understand why the periodic table has as many rows as it does. Each row can roughly be thought of as starting a new energy level. As we proceed across a row, electrons fill energy levels in sections according to where they can be at the lowest energy. So, rather than the row continuing on forever, the periodic table begins a new row which signifies that the electrons in the elements in the next row begin filling a new energy level.



As we saw, sodium doesn't fall to the right of neon on the periodic table just because sodium has more electrons than neon has. Because sodium begins placing its electrons in a new energy level, it is positioned on the far left side at the beginning of a new row.

If we understand a few of the rules about energy level capacity and filling, we can begin to make sense of the periodic table's unusual shape. Why does the first row only consist of two elements? Well, it's because the first energy level can only hold two electrons, and helium, with an atomic number of two, has exactly two electrons. All elements after it have more than 2 electrons, and so they must continue filling their electrons at higher energy levels.

Why does the second period consist of eight elements? It's because the second energy level can only hold eight electrons. If we add in the two the first energy level can hold, the first and second energy levels combined can hold 10 electrons, and neon, the last element of the second period, has exactly 10 electrons.

Though it's a little tricky, potassium begins placing its electrons in the fourth energy level (even with 10 "seats" still available in the third energy level) because the first section of the fourth energy level is lower in energy than the last section of the third energy level.

The number of rows the periodic table has corresponds to the number of energy levels needed to hold all of the electrons of an atom with the greatest known number of electrons.

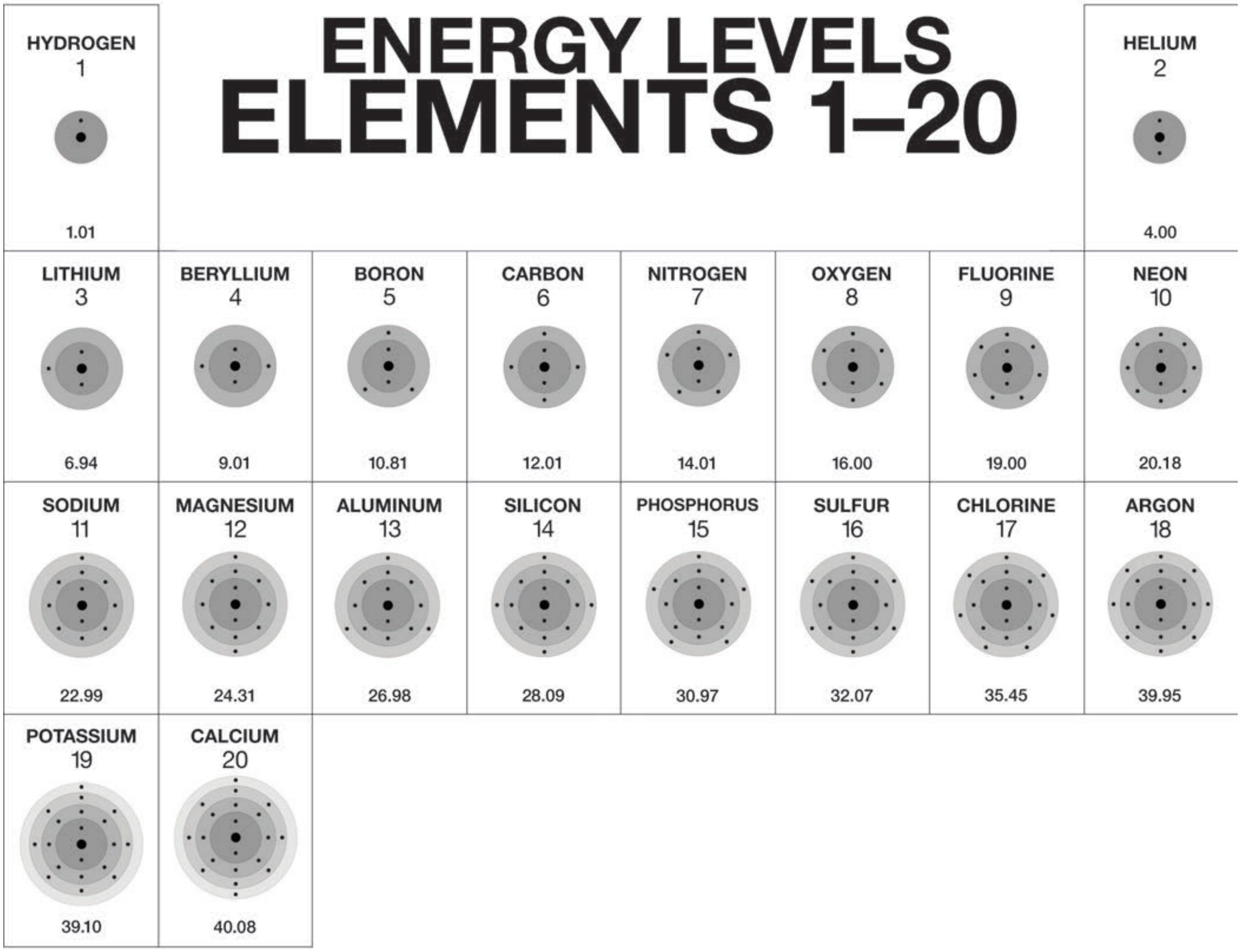
And what about these rows hanging out at the bottom? What's their deal?
The other peculiar feature found in most copies of the periodic table are two mysterious rows, often situated below the rest of the table, which seem to have no relation to the rest of the elements. These rows are called the lanthanide series and actinide series, respectively.

These rows are often placed below the rest of table simply as a convenience. In reality, the elements within the lanthanide series, beginning with the element lanthanum, actually belong alongside barium on the periodic table. Because this would make the table very wide, they are usually placed below the rest of the table so that the format of the periodic table fits more easily on a standard size poster or piece of paper. The same is true for the elements that comprise the actinide series. Beginning with the element actinium, these elements actually belong alongside radium, just below where the lanthanide series would be situated. Some periodic tables actually are formatted in this elongated version. As a convention, however, they are placed below for convenience.

Although alternative forms of the periodic table have been created, some taking unusual shapes like a series of concentric circles in an archery target, the conventional table with the familiar groups and periods is considered the standard.

Why do atoms in the same column have the same number of outer (valence) electrons?
If you think about how the energy levels fill up with electrons, and how the periodic table is designed, you can see how certain atoms end up in the same column. An important point about the columns is that the number of electrons in the outer energy level, called valence electrons, will be the same for all the elements in that column.

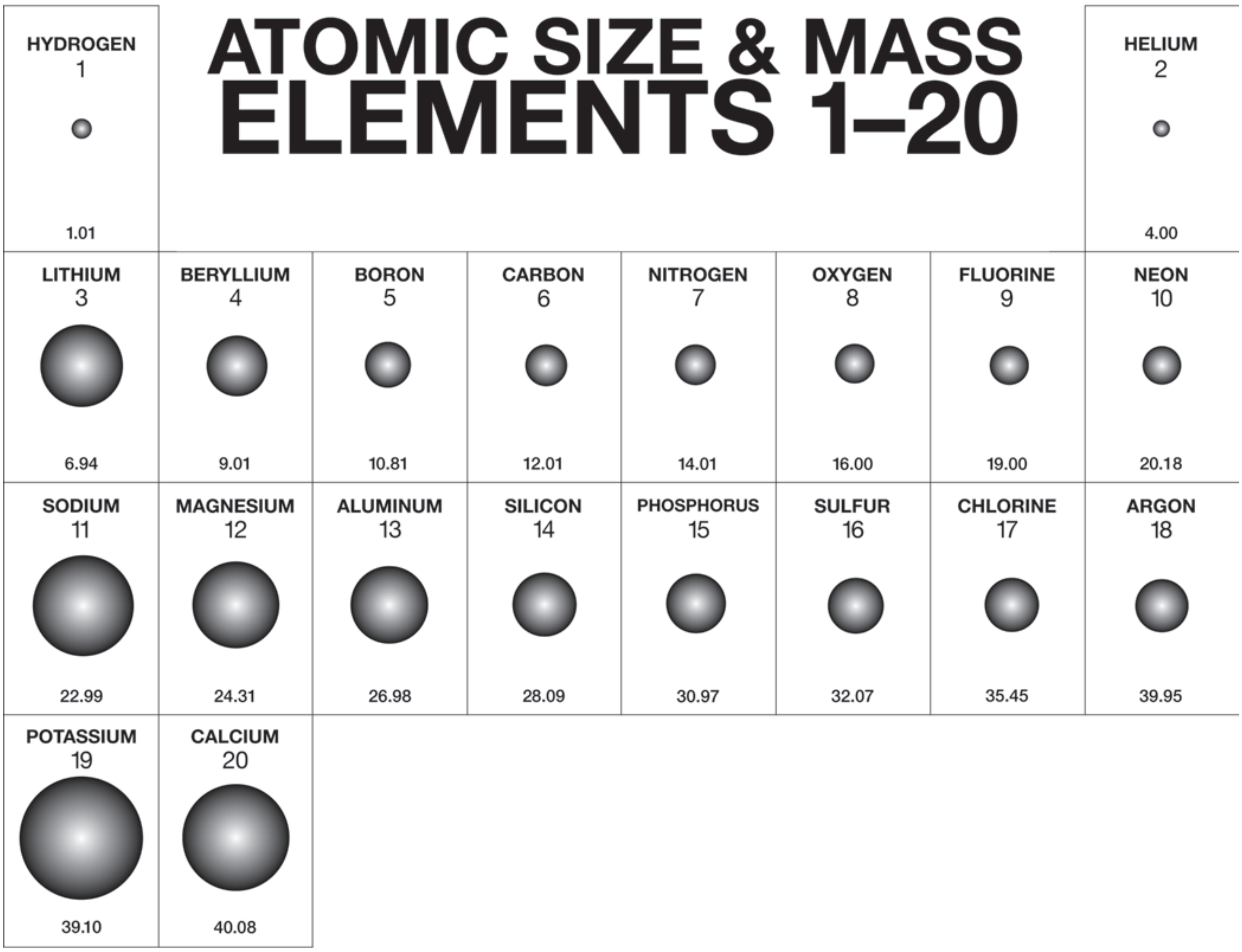
The periodic table is designed so that the first electron starting a new energy level starts a new row on the far left. Each new row starts after the outer energy level of the previous row has eight electrons. An exception to this is starting the second level after the first level has two electrons. Let's look again at the energy level chart.



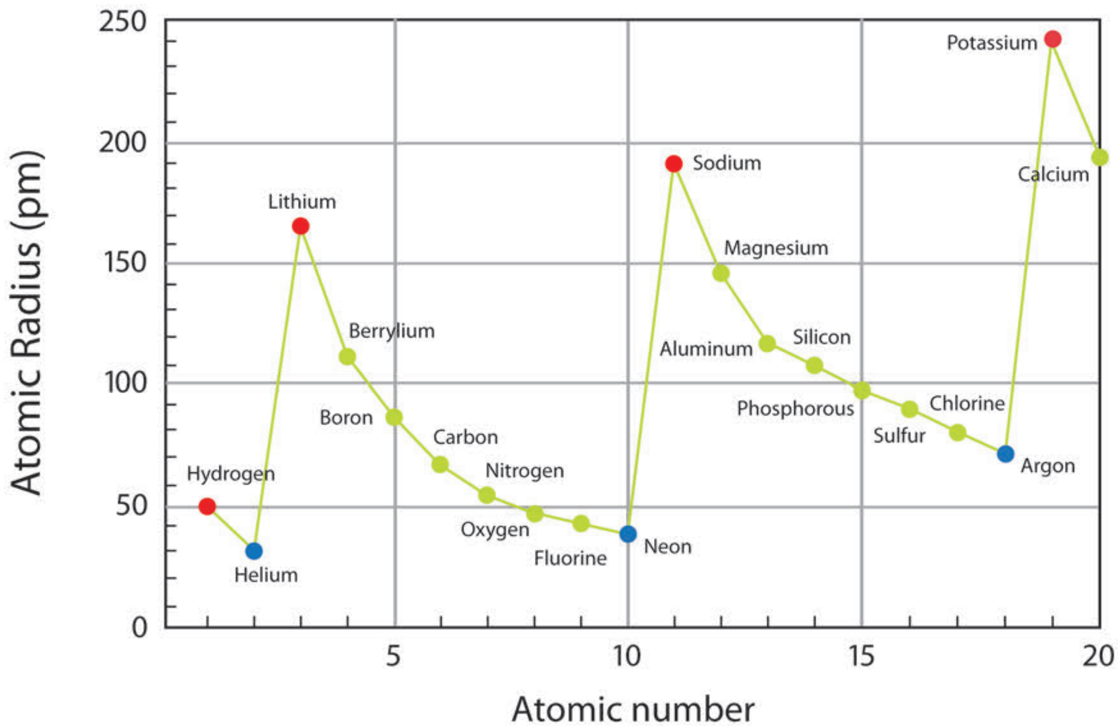
With these principles in mind, you can see why the atoms in the first column, which contains hydrogen (H), lithium (L), sodium (Na), and potassium (K), each have one electron in the outer energy level. In the second column, beryllium (Be), magnesium (Mg) and calcium (Ca), all have two valence electrons. The atoms in the column with boron (B) and aluminum (Al) all have three valence electrons. The atoms in the column with carbon (C) and silicon (Si) have four valence electrons. The rest of the columns follow this same pattern. The transition elements in the middle of the periodic table (not shown in the chart) for the most part have two valence electrons.

What is the significance of the word "periodic" in the periodic table of the elements?
Because of the way the atoms are organized in the periodic table, a pattern of characteristics or properties that repeat "periodically" can be seen from row to row in the table. This is called *periodicity*. Hence the periodic table.

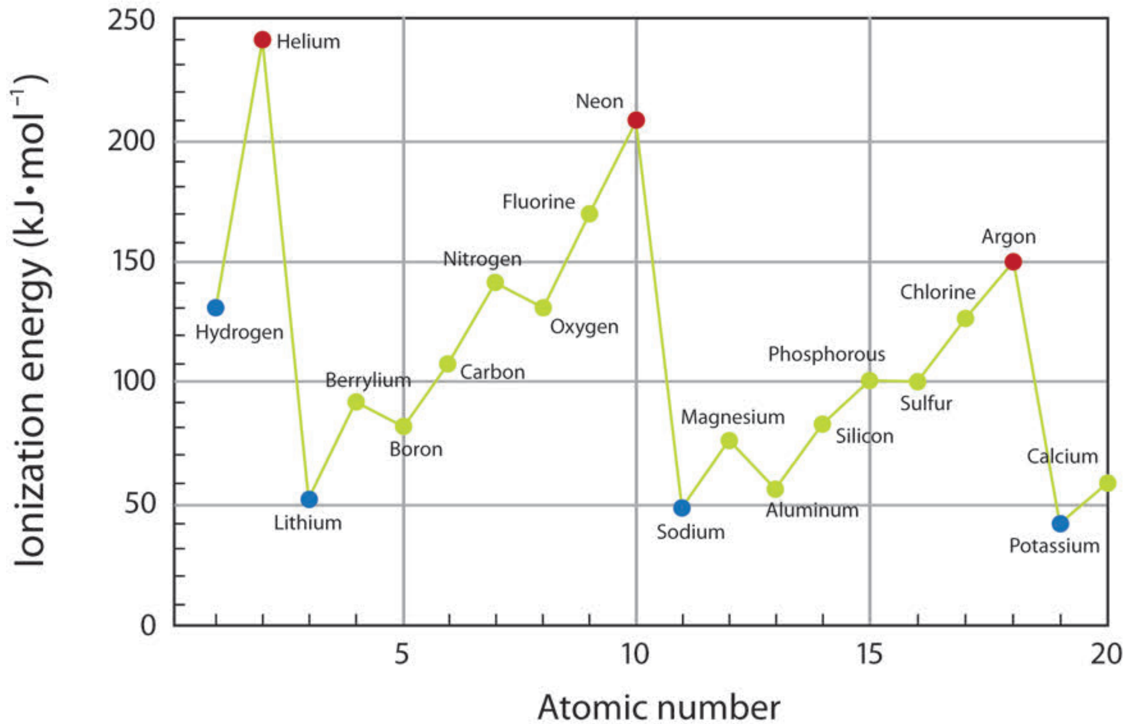
One property that demonstrates the idea of periodicity is atomic radius. Scientists measure atomic radii to tell them how large atoms are. As we proceed across a row (from left to right) we observe that atomic radii decrease. For example, magnesium has a smaller atomic radius than sodium, and aluminum has a smaller atomic radius than magnesium, and so on. This same pattern repeats itself in the next row and the next row in a periodic way.



***Note**: Some charts of atomic radii show the atoms in the last column on the right (Helium, Neon, Argon …) having larger radii than one or more of the atoms to their left. This is a result of using a different measuring technique. For our purposes, the trend is decreasing atomic size as you move from left to right along a row.*



Another example of periodicity is a property called *ionization energy*. Ionization energy refers to the amount of energy needed to remove an electron from an atom to form an ion. The more difficult it is to remove an electron from an atom, the higher its ionization energy. As a trend, ionization energy increases as you move across a row (from left to right). For example, in the first row, Hydrogen, on the far left, has a low ionization energy and Helium, on the far right, has a high ionization energy. Each row begins with a low value and ends with a high value.



There are other periodic trends in the properties of atoms. You can see some of these by going to *http://acswebcontent.acs.org/games/pt.html*.

Click the “Plot Data” tab to get to a drop-down menu of different properties you can choose. The first one displayed is “Molar Mass” which shows the mass of a mole of each atom. This can be thought of as comparing the atomic mass of one atom to another. This property is not periodic. You can choose atomic radius or ionization energy to see the trends discussed earlier or select another property.

Chapter 4, Lesson 4:
Energy Levels, Electrons, and Covalent Bonding

Key Concepts

- The electrons on the outermost energy level of the atom are called valence electrons.
- The valence electrons are involved in bonding one atom to another.
- The attraction of each atom's nucleus for the valence electrons of the other atom pulls the atoms together.
- As the attractions bring the atoms together, electrons from each atom are attracted to the nucleus of both atoms, which "share" the electrons.
- The sharing of electrons between atoms is called a covalent bond, which holds the atoms together as a molecule.
- A covalent bond happens if the attractions are strong enough in both atoms and if each atom has room for an electron in its outer energy level.
- Atoms will covalently bond until their outer energy level is full.
- Atoms covalently bonded as a molecule are more stable than they were as separate atoms.

Summary

Students will look at animations and refer to the energy level models they have been using to make drawings of the process of covalent bonding. Students will consider why atoms bond to form molecules like H_2 (hydrogen), H_2O (water), O_2 (oxygen), CH_4 (methane), and CO_2 (carbon dioxide).

Objective

Students will be able to explain that attraction between the protons and electrons of two atoms cause them to bond. Students will be able to draw a model of the covalent bonds between the atoms in H_2 (hydrogen), H_2O (water), O_2 (oxygen), CH_4 (methane), and CO_2 (carbon dioxide).

Evaluation

The activity sheet will serve as the "Evaluate" component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding.

Safety

Be sure you and the students wear properly fitting goggles.

Materials for Each Group

- 9-volt battery
- 2 wires with alligator clips on both ends
- 2 pencils sharpened at both ends
- Water
- Epsom Salt (magnesium sulfate)
- Clear plastic cup
- Tape

About this Lesson

This lesson will probably take more than one class period.

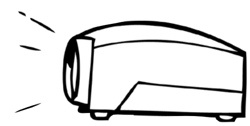
ENGAGE

1. Show an animation to introduce the process of covalent bonding.

Introduce the question students will investigate in this lesson:

- If atoms have an equal number of protons and electrons, why do atoms bond to other atoms? Why don't they just stay separate?
  Begin to answer this question by using hydrogen as an example.

Project the animation *Covalent bond in hydrogen.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson4#covalent_bond_hydrogen_animation



Make sure students see that each hydrogen atom has 1 proton and 1 electron. Remind students that the electron and its own proton are attracted to each other. Explain that if the atoms get close enough to each other, the electron from each hydrogen atom feels the attraction from the proton of the other hydrogen atom (shown by the double-headed arrow). Point out to students that the attractions are not strong enough to pull the electron completely away from its own proton. But the attractions are strong enough to pull the two atoms close enough together so that the electrons feel the attraction from both protons and are shared by both atoms. At the end of the animation, explain that the individual hydrogen atoms have now bonded to become the molecule H_2. This type of bond is called a *covalent* bond. In a covalent bond, electrons from each atom are attracted or "shared" by both atoms.

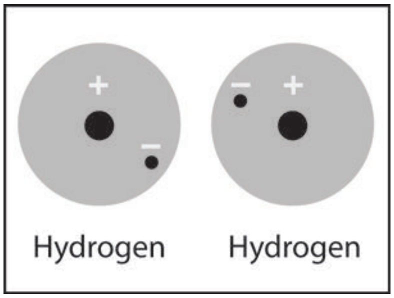
EXPLAIN

2. Discuss the conditions needed for covalent bonding and the stable molecule that is formed.

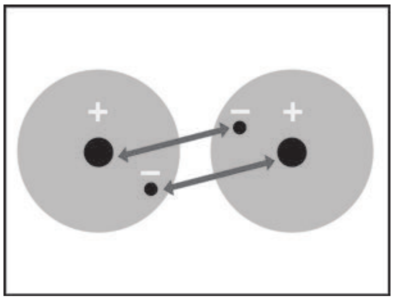
Project the image *Covalent bond in hydrogen.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson4#covalent_bond_hydrogen_illustrations

Note: *This model of covalent bonding for the hydrogen molecule (H_2) starts with 2 individual hydrogen atoms. In reality, hydrogen atoms are never separate to start with. They are always bonded with something else. To simplify the process, this model does not show the hydrogen atoms breaking their bonds from other atoms. It only focuses on the process of forming covalent bonds between two hydrogen atoms.*

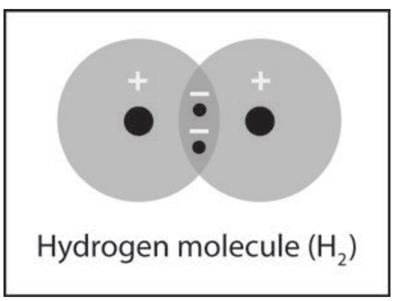
Read more about bonding in the additional teacher background section at the end of this lesson.



Two hydrogen atoms are near each other.



When two hydrogen atoms come close enough to each other, their electrons are attracted to the proton of the other atom.



Because there is both a strong enough attraction between atoms and room for electrons in the outer energy level of both atoms, the atoms share electrons. This forms a covalent bond.

Tell students that there are two main reasons why two hydrogen atoms bond together to make one hydrogen molecule:

- There needs to be a strong enough attraction between the electrons of each atom for the protons of the other atom.
- There needs to be room in the outer energy level of both atoms.

Once bonded, the hydrogen molecule is more stable than the individual hydrogen atoms. Explain to students that by being part of a covalent bond, the electron from each hydrogen atom gets to be near two protons instead of only the one proton it started with. Since the electrons are closer to more protons, the molecule of two bonded hydrogen atoms is more stable than the two individual unbonded hydrogen atoms.

This is why it is very rare to find a hydrogen atom that is not bonded to other atoms. Hydrogen atoms bond with other hydrogen atoms to make hydrogen gas (H_2). Or they can bond with other atoms like oxygen to make water (H_2O) or carbon to make methane (CH_4) or many other atoms.

3. Show students that when two hydrogen atoms bond together, the outer energy level becomes full.

Have students look at their *Periodic table of energy levels for elements 1–20* distributed in lesson 3.

Explain that the two electrons in the hydrogen molecule (H_2) can be thought of as "belonging" to each atom. This means that each hydrogen atom now has two electrons in its first energy level. The first energy level in the outer energy level for hydrogen and can only accommodate or "hold" two electrons. Atoms will continue to covalently bond until their outer energy levels are full. At this point, additional atoms will not covalently bond to the atoms in the H_2 molecule.

4. Have students describe covalent bonding in a hydrogen molecule on their activity sheet and then review their answers.

Give each student an activity sheet.
Have students write a short caption under each picture to describe the process of covalent bonding and answer the first three questions. The rest of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions.

Ask students:

- **What did you write for the second and third pictures of covalent bonding?**
  Center drawing: When two hydrogen atoms come close enough, their electrons are attracted to the proton of the other atom.
  Last drawing: This brings the atoms close enough together that they share electrons.
- **What are two conditions atoms must have in order to form covalent bonds with one another?**
  There is a strong enough attraction between atoms and there is room for electrons in the outer energy level of both atoms.
- **Why is a hydrogen molecule (H_2) more stable than two individual hydrogen atoms?**
  In the hydrogen molecule, the electrons from each atom are able to be near two protons instead of only the one proton it started with. Whenever negative electrons are near additional positive protons, the arrangement is more stable.
- **Why doesn't a third hydrogen atom join the H_2 molecule to make H_3?**
  When two hydrogen atoms share their electrons with each other, their outer energy levels are full.

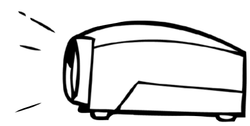
You could explain to students that when the outer energy levels are full, sharing electrons with another atom would not happen for two main reasons:

1. An electron from a new atom would have to join an atom in the H_2 molecule on the next energy level, further from the nucleus where it would not feel a strong enough attraction.
2. An electron from an atom already in the H_2 molecule and close to the nucleus would need to move further away to share with the new atom.

Both of these possibilities would make the molecule less stable and therefore would not happen.

5. Discuss the process of covalent bonding in a water molecule.

Project the animation *Covalent bond in water*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson4#covalent_bonding_water



Before hitting the "play" button, point out the oxygen atom and the two hydrogen atoms.

Ask students:

- **Is there anything that might attract these atoms to one another?**
  Students should suggest that the electrons from each atom are attracted to the protons of the other atoms.

Play the animation to show the attraction between the protons of oxygen for the electron from each of the hydrogen atoms, the attraction of the proton from the hydrogen atoms for the electrons of oxygen, and the atoms coming together.

Explain that the electrons are shared by the oxygen and hydrogen atoms forming a covalent bond. These bonds hold the oxygen and hydrogen atoms together and form the H_2O molecule. The reason why the atoms are able to bond is that the attractions are strong enough in both directions and there is room for the electrons on the outer energy level of the atoms.

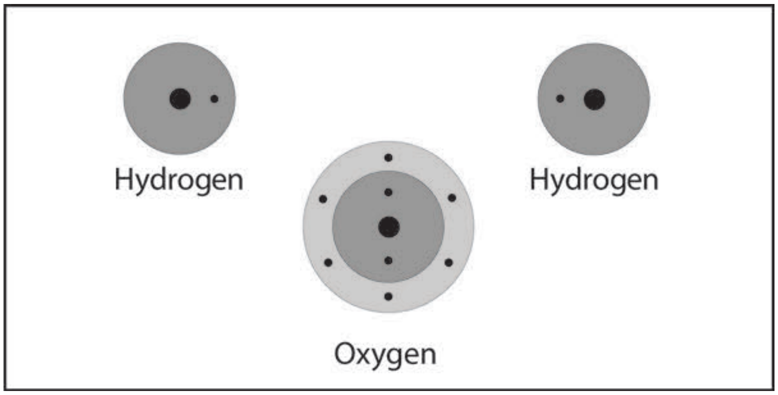
The electron from each hydrogen atom and the electrons from the oxygen atom get to be near more protons when the atoms are bonded together as a molecule than when they are separated as individual atoms. This makes the molecule of bonded oxygen and hydrogen atoms more stable than the individual separated atoms.

Explain to students that the two electrons in the bond between the hydrogen atom and the oxygen atom can be thought of as "belonging" to each atom. This gives each hydrogen atom two electrons in its outer energy level, which is full. It also gives oxygen 8 electrons in its outer energy level, which is also full.

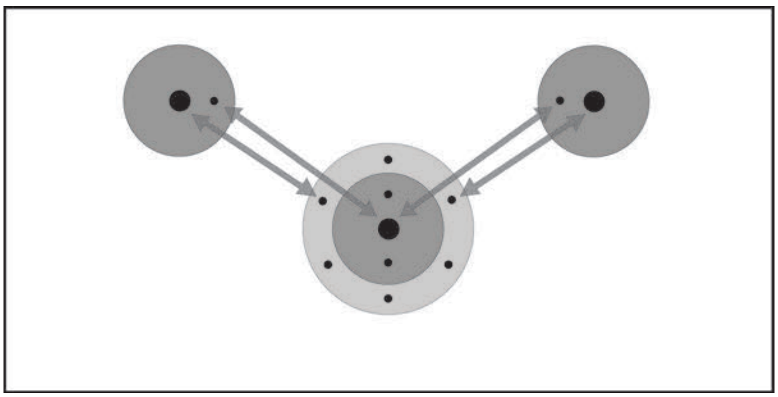
Project the image *Covalent bond in water*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson4#covalent_bonding_water_illustrations
Review with students the process of covalent bonding covered in the animation.

6. Have students describe covalent bonding in a water molecule on their activity sheet.

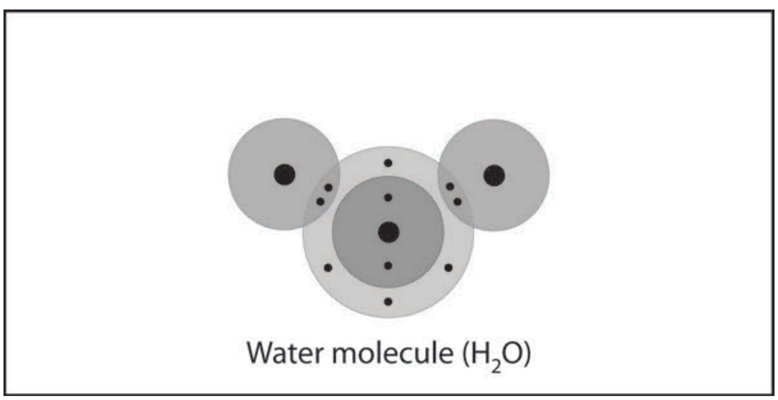
Have students write a short caption beside each picture to describe the process of covalent bonding in the water molecule.



Two hydrogen atoms and one oxygen atom are near each other.



When two hydrogen atoms come close enough to an oxygen atom, their electrons are attracted to the proton of the other atom.



Because there is both a strong enough attraction between atoms and room for electrons in the outer energy levels of the atoms, they share electrons. This forms a covalent bond.

***Note**: This model of covalent bonding for a water molecule starts with 2 individual hydrogen atoms and 1 oxygen atom. In reality, these atoms are never separate to start with. They are always bonded with something else. To simplify the process, this model does not show the hydrogen and oxygen atoms breaking their bonds from other atoms. It only focuses on the process of forming covalent bonds to make water.*

Ask students:

- **Why can't a third hydrogen atom join the water molecule (H_2O) to make H_3O?**
  Once the outer energy levels are full, sharing electrons with another atom would not happen for two main reasons: An electron from a new atom would have to join an atom in the H_2O molecule on the next energy level, further from the nucleus where it would not feel a strong enough attraction. An electron from an atom already in the H_2O molecule and close to the nucleus would need to move further away to share with the new atom. Both of these possibilities would make the molecule less stable and would not happen.

EXPLORE

7. Have students use electricity to break the covalent bonds in water molecules.

Tell students that electrical energy can be used to break the covalent bonds in water molecules to produce hydrogen atoms and oxygen atoms. Two hydrogen atoms then bond to form hydrogen gas (H_2) and two oxygen atoms bond to form oxygen gas (O_2).

You may choose to do this activity as a demonstration or show the video *Electrolysis*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson4#electrolysis

Question to investigate
What is produced when the covalent bond in water molecules is broken?

Materials for each group
- 9-volt battery
- 2 wires with alligator clips on both ends
- 2 pencils sharpened at both ends
- Water
- Epsom salt (magnesium sulfate)
- Clear plastic cup
- Tape

Procedure
1. Place a battery between 2 pencils. Be sure that the battery is more than half-way up.
2. With the help of a partner, wrap tape around the pencils and battery as shown.
3. Add water to a clear plastic cup until it is about ½-full.
4. Add about ½ teaspoon of Epsom salt to the water and stir until the salt dissolves.
5. Connect one alligator clip to one terminal of the battery.
6. Using the other wire, connect one alligator clip to the other terminal of the battery.

7. Connect one end of the pencil lead to the alligator clip at the end of one of the wires.
8. Using the other wire, connect one end of the other pencil lead to the alligator clip at the end of the wire.
9. Place the ends of the pencil into the water as shown.

Expected results
Bubbles will form and rise initially from one pencil lead. Soon, bubbles will form and rise from the other. Students should be able to see that there is more of one gas than the other. The gas that forms the small bubbles that comes off first is hydrogen. The other gas that forms the larger bubbles and lags behind a bit is oxygen.

***Note**: There will be bubbling when hydrogen and oxygen gas form on the pencil leads. Be sure students do not get the misconception that the bubbles they see mean that the water is boiling. In boiling, the bonds holding the atoms together in water molecules do not come apart. In the process of electrolysis, the bonds holding the atoms together do come apart.*

8. Discuss student observations.

Ask students:

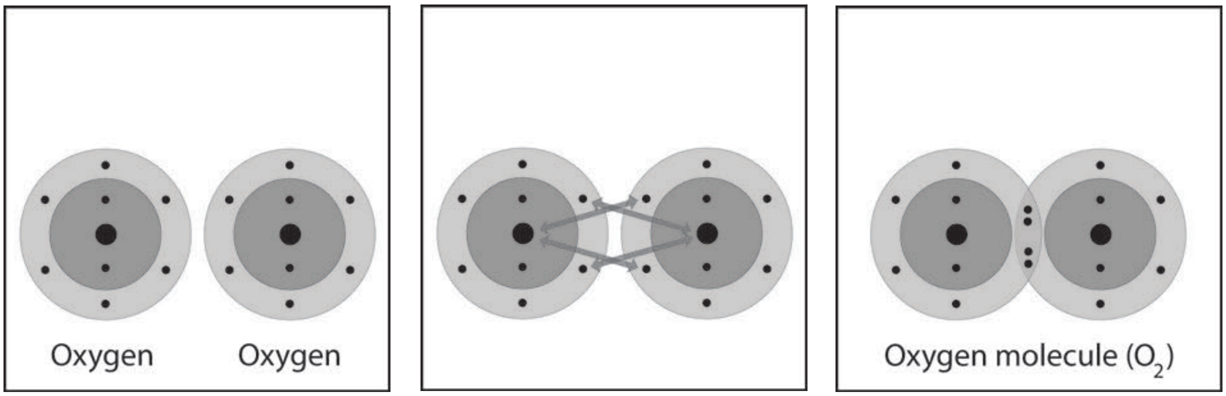
- **What are the bubbles made out of in the activity?**
  Hydrogen gas (H_2) and oxygen gas (O_2)
- **Why was there more hydrogen gas produced than oxygen gas?**
  Each water molecule breaks into 2 hydrogen atoms and 1 oxygen atom. Two hydrogen atoms then bond to form hydrogen gas (H_2) and 2 oxygen atoms bond to form oxygen gas (O_2). Each water molecule has all the atoms needed to make 1 molecule of hydrogen gas. But with only 1 oxygen atom, a water molecule only has half of what is needed to make 1 molecule of oxygen gas. So, 2 water molecules will produce 2 molecules of hydrogen gas but only 1 molecule of oxygen gas.

EXTEND

9. Help students understand how atoms combine to form the molecules of oxygen, methane, and carbon dioxide.

Remind students that in this lesson they looked at the covalent bonds in hydrogen molecules and in water molecules. Tell them that they will look at the covalent bonds in three other common substances.

Project the animation *Oxygen's double bond*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson4#oxygen_double_bond
Explain to students that the oxygen molecules that are present in our air are made up of 2 oxygen atoms. This animation will show them what the covalent bond between 2 oxygen atoms is like. Narrate the animation by pointing out that each oxygen atom has 6 valence electrons. When the oxygen atoms get close together, the attractions from the nucleus of both atoms attract the outer electrons. In this case, 2 electrons from each atom are shared. This is called a double bond.



Each oxygen atom has 6 valence electrons in its outer energy level.

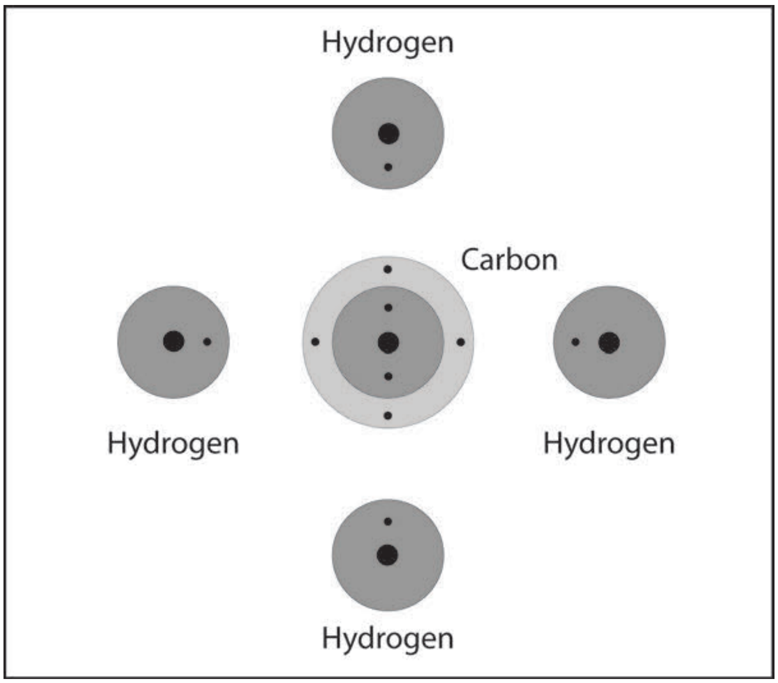
When two oxygen atoms get close to each other, the attractions from the nucleus of both atoms attract the outer electrons.

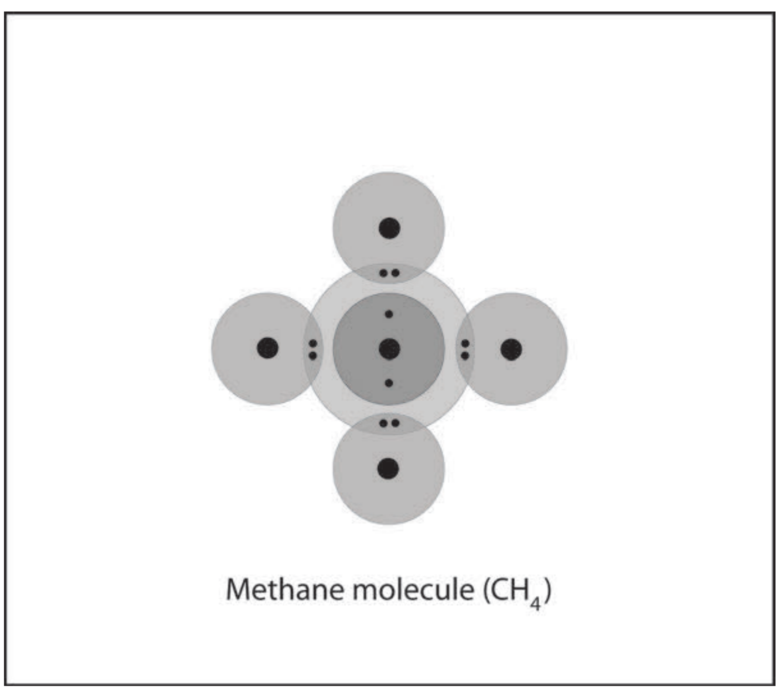
In this case, *two* electrons from each atoms are shared. This is called a double bond.

Project the image *Oxygen's double bond II*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson4#oxygen_double_bond_illustrations
Review with students the process of covalent bonding covered in the animation.

Project the before and after pictures *Covalent bonding of methane*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson4#covalent_bonding_methane
Ask students:

- **Briefly describe the process of covalent bonding between the carbon and the four hydrogen atoms to make a methane molecule. Be sure to mention attractions between electrons and protons and the number of electrons in the outer energy level for the atoms in the final molecule.**
  Be sure students realize that the protons of each atom attracts the other atoms electrons, which brings the atoms together. Atoms continue to bond with other atoms until their outer energy levels are full.

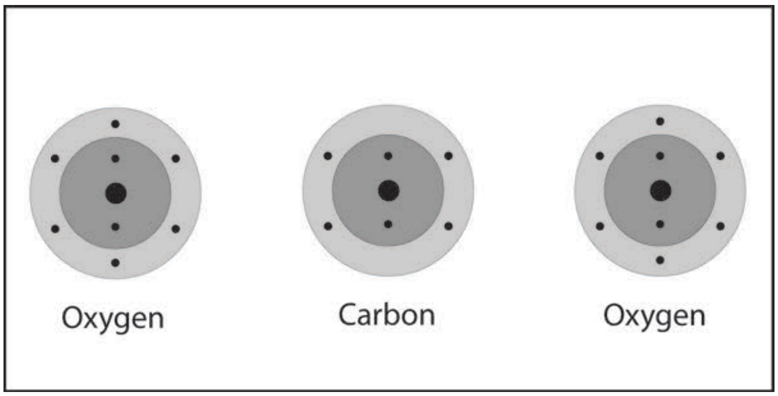


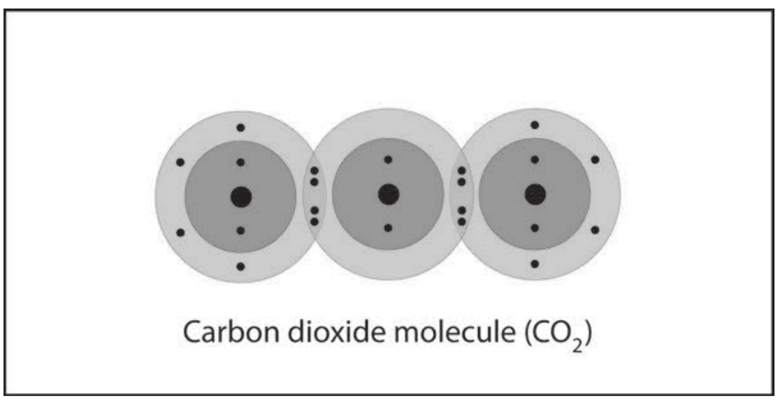


Project the before and after pictures *Covalent bonding of carbon dioxide gas*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson4#covalent_bond_carbon_dioxide

Ask students:

- **Briefly describe the process of covalent bonding between the carbon and the two oxygen atoms to make a carbon dioxide molecule. Be sure to mention attractions between electrons and protons and the number of electrons in the outer energy level for the atoms in the final molecule.**
  Be sure students realize that the protons of each atom attracts the other atoms electrons, which brings the atoms together. Atoms continue to bond with other atoms until their outer energy levels are full.





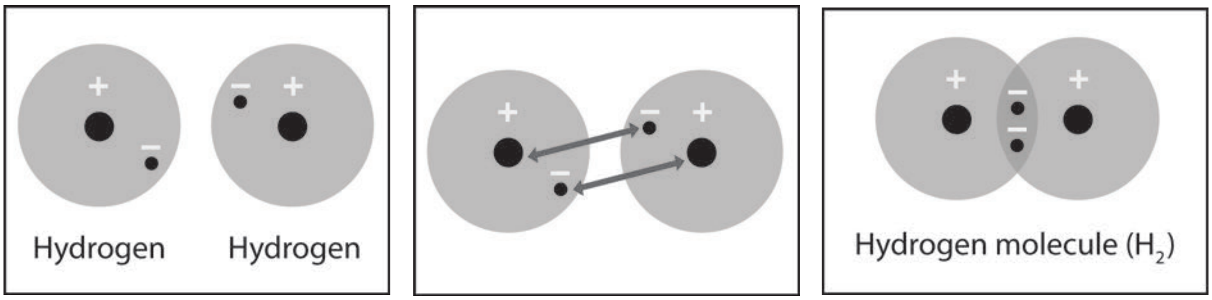
Activity Sheet
Chapter 4, Lesson 4
Energy Levels, Electrons, and Covalent Bonding

Name ____________________

Date _____________________

EXPLAIN IT WITH ATOMS & MOLECULES

1. Write a short caption under each picture to describe the process of covalent bonding.



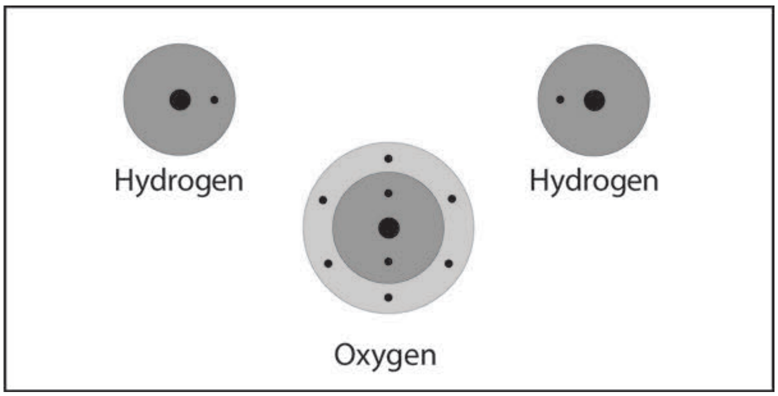
Two hydrogen atoms are near each other.

2. What are two conditions atoms must have in order to form covalent bonds with one another?

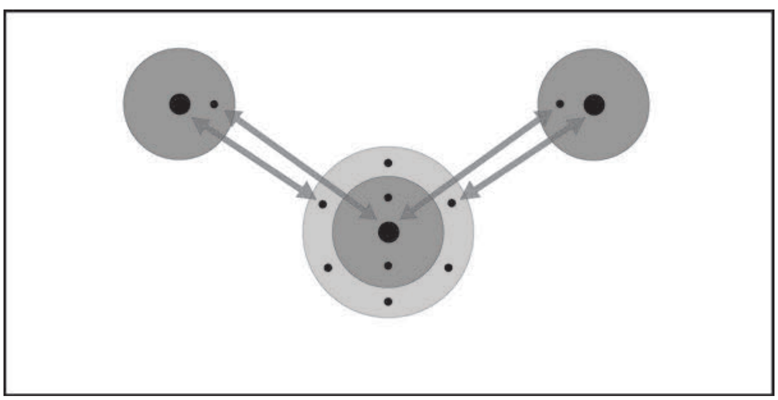
3. Why is a hydrogen molecule (H_2) more stable than two individual hydrogen atoms?

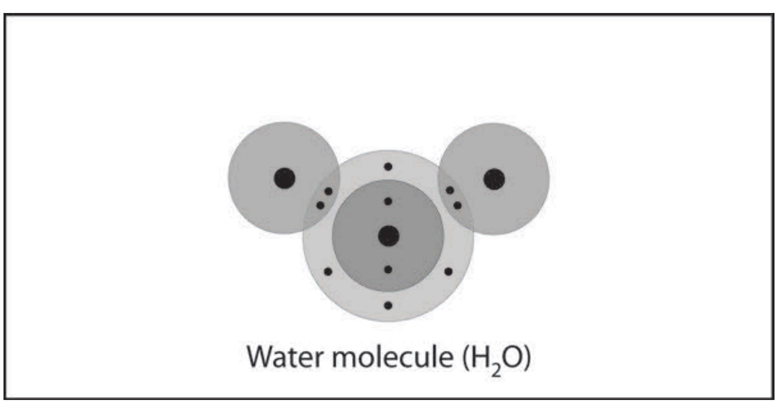
4. Why can't a third hydrogen atom join the H_2 molecule to make H_3?

5. Write a short caption beside each picture to describe the process of covalent bonding.



Two hydrogen atoms and one oxygen atom are near each other.





6. Why can’t a third hydrogen atom join the water molecule (H_20) to make H_3O?

ACTIVITY

Question to investigate
What is produced when the covalent bond in water molecules is broken?

Materials for each group

- 9-volt battery
- 2 wires with alligator clips on both ends
- 2 pencils sharpened at both ends
- Water
- Epsom salt (magnesium sulfate)
- Clear plastic cup
- Tape

Procedure

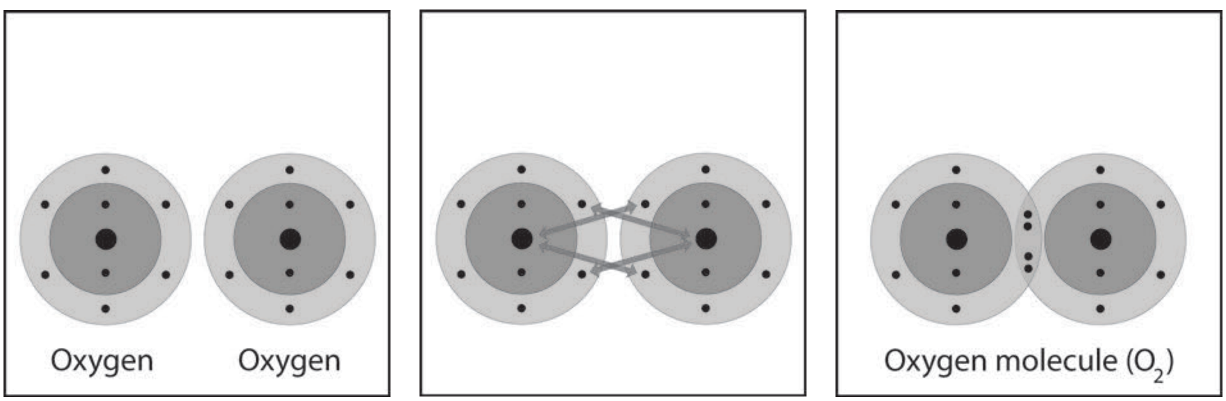
1. Place a battery between 2 pencils. Be sure that the battery is more than half-way up.
2. With the help of a partner, wrap tape around the pencils and battery as shown.
3. Add water to a clear plastic cup until it is about ½-full.
4. Add about a ½ teaspoon of Epsom salt to the water and stir until the salt dissolves.
5. Connect one alligator clip to one terminal of the battery.
6. Using the other wire, connect one alligator clip to the other terminal of the battery.
7. Connect one end of the pencil lead to the alligator clip at the end of one of the wires.
8. Using the other wire, connect one end of the other pencil lead to the alligator clip at the end of the wire.
9. Place the ends of the pencil into the water as shown.

7. What were the bubbles made out of in this activity?

8. Why was there more hydrogen gas produced than oxygen gas?
   HINT: Look back at the drawings showing the number of hydrogen and oxygen atoms that bond to form a water molecule.

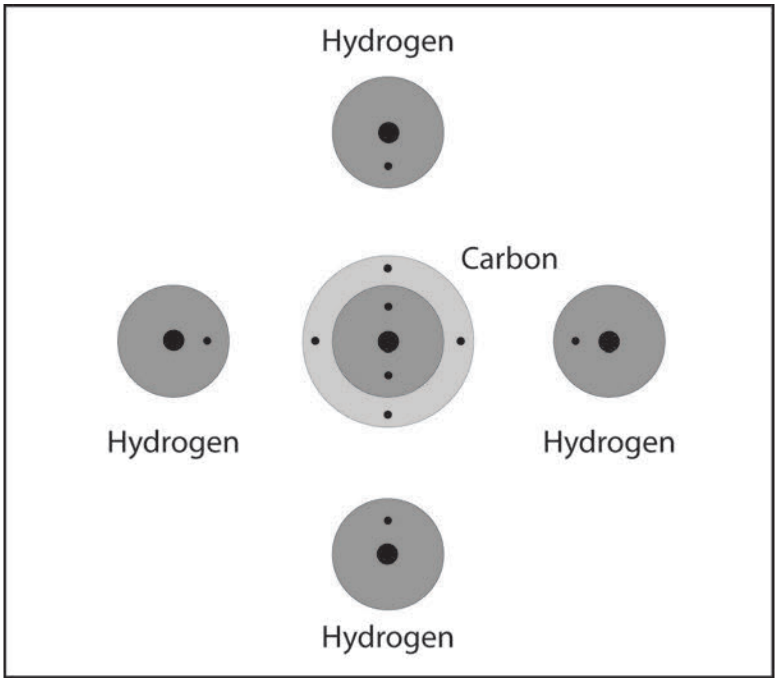
TAKE IT FURTHER

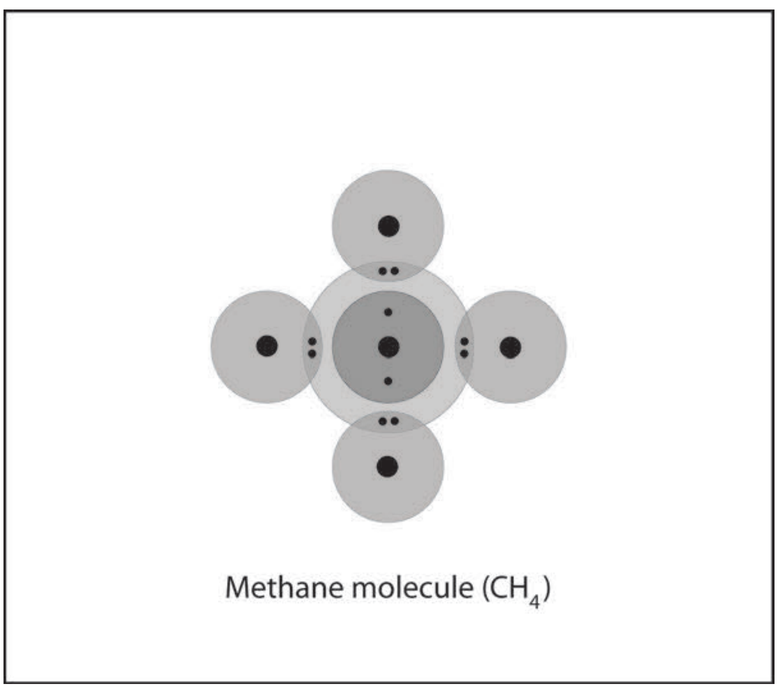
9. Briefly describe the process of covalent bonding between two oxygen atoms to make an oxygen molecule. Be sure to mention attractions between electrons and protons and the number of electrons in the outer energy level for the atoms in the final molecule.



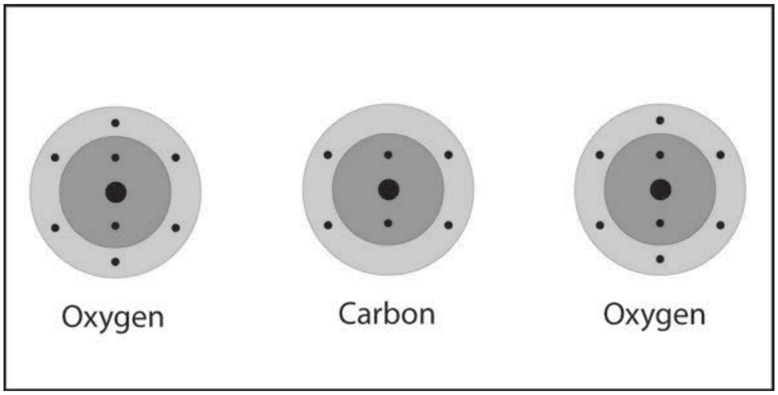
Each oxygen atom has 6 valence electrons in its outer energy level.

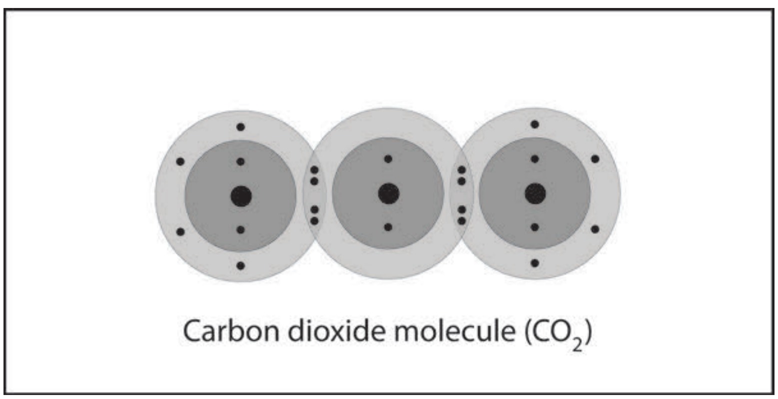
10. Briefly describe the process of covalent bonding between the carbon and the four hydrogen atoms to make a methane molecule. Be sure to mention attractions between electrons and protons and the number of electrons in the outer energy level for the atoms in the final molecule.





11. Briefly describe the process of covalent bonding between the carbon and the two oxygen atoms to make a carbon dioxide molecule. Be sure to mention attractions between electrons and protons and the number of electrons in the outer energy level for the atoms in the final molecule.





Additional Teacher Background
Chapter 4, Lesson 4, p. 318

A common approach to figuring out how atoms bond covalently and ionically is to use the “octet rule”. This rule relies on the fact that atoms bond until they have 8 electrons in their outer energy levels or 2 electrons in the outer level in the case of hydrogen and helium. It is often stated that atoms “want” to have 8 electrons in their outer energy level so they bond until they have 8 as if having 8 electrons is a goal in itself.

The approach taken in Lesson 4 and 5 achieves the same result but it does not use the goal of having 8 electrons or wanting 8 electrons as the reason why atoms bond. Instead the approach emphasizes the fact that, if the attractions are favorable in both directions and there is room to accommodate electrons, atoms continue to bond until it is unfavorable to do so. This occurs when the outer energy levels of the atoms are full.

Chapter 4, Lesson 5: Energy Levels, Electrons, and Ionic Bonding

Key Concepts

- The attractions between the protons and electrons of atoms can cause an electron to move completely from one atom to the other.
- When an atom loses or gains an electron, it is called an ion.
- The atom that loses an electron becomes a positive ion.
- The atom that gains an electron becomes a negative ion.
- A positive and negative ion attract each other and form an ionic bond.

Summary

Students will look at animations and make drawings of the ionic bonding of sodium chloride ($NaCl$). Students will see that both ionic and covalent bonding start with the attractions of protons and electrons between different atoms. But in ionic bonding, electrons are transferred from one atom to the other and not shared like in covalent bonding. Students will use Styrofoam balls to make models of the ionic bonding in sodium chloride (salt).

Objective

Students will be able to explain the process of the formation of ions and ionic bonds.

Evaluation

The activity sheet will serve as the "Evaluate" component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles.

Materials for Each Group

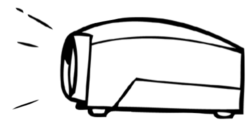
- Black paper
- Salt
- Cup with salt from evaporated saltwater
- Magnifier
- Permanent marker

Materials for Each Student

- 2 small Styrofoam balls
- 2 large Styrofoam balls
- 2 toothpicks

Note: In an ionically bonded substance such as NaCl, the smallest ratio of positive and negative ions bonded together is called a "formula unit" rather than a "molecule." Technically speaking, the term "molecule" refers to two or more atoms that are bonded together covalently, not ionically. For simplicity, you might want to use the term "molecule" for both covalently and ionically bonded substances.

ENGAGE



1. Show a video of sodium metal reacting with chlorine gas.

Project the video *Sodium and chlorine react.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson5#sodium_chlorine_react
Before starting the video, tell students that chlorine is a greenish poisonous gas and sodium is a shiny, soft, and very reactive metal. But when they react, they form sodium chloride (table salt). Tell students that in the video, the drop of water helps expose the atoms at the surface of the sodium so that they can react with the chlorine. The formation of the salt crystals releases a lot of energy.

Note: If students ask if the salt they eat is made this way in salt factories, the answer is no. The salt on Earth was produced billions of years ago but probably not from pure chlorine gas and sodium metal. These days, we get salt from mining it from a mineral called halite or from evaporating sea water.

EXPLAIN

2. Show an animation to introduce the process of ionic bonding.

Project the animation *Ionic bond in sodium chloride.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson5#ionic_bond_in_sodium_chloride
Remind students that in covalent bonding, atoms share electrons. But there is another type of bonding where atoms don't share, but instead either take or give up electrons. This is called ionic bonding. This animation shows a very simplified model of how sodium and chloride ions are formed.

Note: In order to simplify the model of ionic bonding, a single atom of sodium and chlorine are shown. In reality, the chlorine atom would be bonded to another chlorine atom as part of the gas Cl_2. The sodium atom would be one of billions of trillions of sodium atoms bonded together as a solid. The combination of these substances is a complex reaction between the atoms of the two substances. The animation shows single separated atoms to illustrate the idea of how ions and ionic bonds are formed.

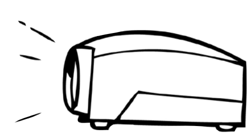
Explain what happens during the animation.
Tell students that the attraction of the protons in the sodium and chlorine for the other atom's electrons brings the atoms closer together. Chlorine has a stronger attraction for electrons than sodium (shown by the thicker arrow). At some point during this process, an electron from the sodium is transferred to the chlorine. The sodium loses an electron and the chlorine gains an electron.

Tell students that when an atom gains or loses an electron, it becomes an *ion*.
- Sodium loses an electron, leaving it with 11 protons, but only 10 electrons. Since it has 1 more proton than electrons, sodium has a charge of +1, making it a positive ion.
- Chlorine gains an electron, leaving it with 17 protons and 18 electrons. Since it has 1 more electron than protons, chlorine has a charge of –1, making it a negative ion.
- When ions form, atoms gain or lose electrons until their outer energy level is full.
  - For example, when sodium loses its one outer electron from the third energy level, the second level becomes the new outer energy level and is full. Since these electrons are closer to the nucleus, they are more tightly held and will not leave.
  - When chlorine gains an electron, its third energy level becomes full. An additional electron cannot join, because it would need to come in at the fourth energy level. This far from the nucleus, the electron would not feel enough attraction from the protons to be stable.
- Then the positive sodium ion and negative chloride ion attract each other and form an ionic bond. The ions are more stable when they are bonded than they were as individual atoms.

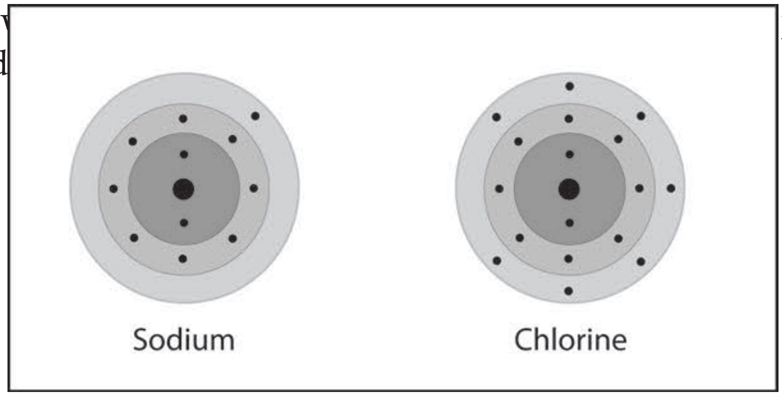
3. Have students describe the process of ionic bonding in sodium chloride on their activity sheet.

Give each student an activity sheet.
Have students write a short caption under each picture to describe the process of covalent bonding and answer the first three questions. The rest of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions.

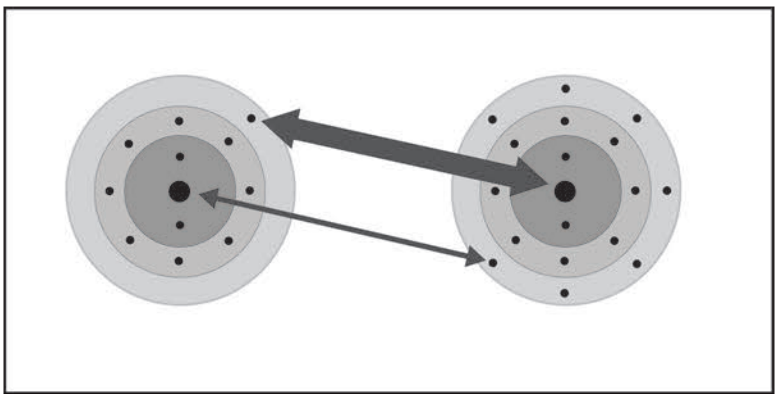
Project the image *Ionic bond in sodium chloride*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson5# ionic_bond_in_sodium_chloride_2
Review with students the process of ionic bonding covered in the animation.



Help students v… process of ionic bonding in sod…

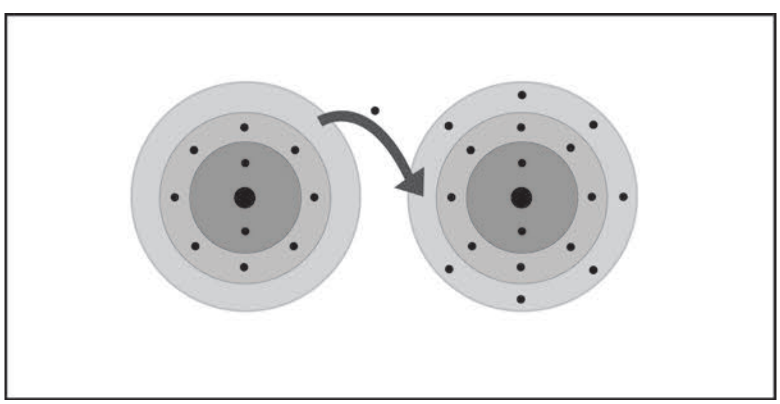


Sodium and chlorine atoms are near each other.

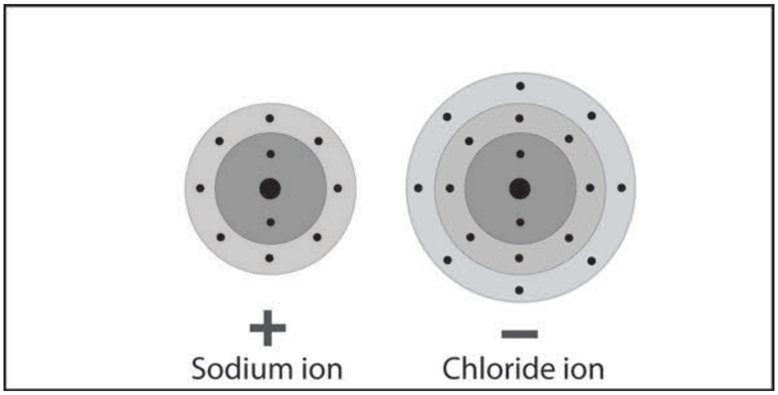


The protons of the two atoms attact the electrons of the other atom.

The thicker arrow shows that chlorine has a stronger attraction for electrons than sodium has.

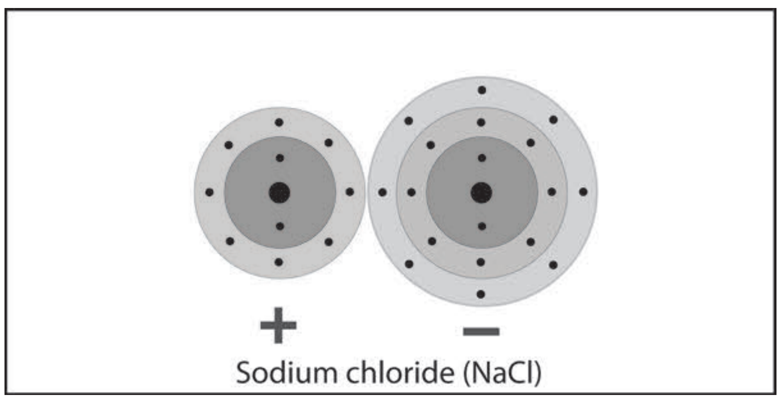


During the interactions between the atoms, the electron in sodium's outer energy level is transferred to the outer energy level of the chlorine atom.

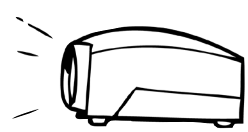


Since sodium *lost* an electron, it has 11 protons, but only 10 electrons. This makes sodium a *positive* ion with a charge of +1.

Since chlorine *gained* an electron it has 17 protons and 18 electrons. This makes chloride a *negative* ion with a charge of –1.



The positive sodium ion and negative chloride ion attract one another. They make an ionic bond and form the ionic compound NaCl.



4. Show students a model of a sodium chloride crystal and have them identify the ions.

Project the image *Sodium chloride crystal*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson5#sodium_chloride_crystal

Review with students the process of ionic bonding covered in the animation so that students will understand why the sodium ions are positive and the chloride ions are negative. Remind students that the scale of any model of atoms, ions, or molecules is enormous compared to the actual size. In a single grain of salt there are billions of trillions of sodium and chloride ions.

Ask students:

- **What ion is the larger ball with the negative charge?**
  The chlorine ion.
- **What made it negative?**
  It gained an electron.
- **What is the ion with the positive charge?**
  The sodium ion.
- **What made it positive?**
  It lost an electron.

EXPLORE

5. Have students observe actual sodium chloride crystals and relate their shape to the molecular model.

This two-part activity will help students see the relationship between the arrangement of ions in a model of a sodium chloride crystal and the cubic shape of real sodium chloride crystals.

Teacher preparation
The day before the lesson, dissolve about 10 grams of salt in 50 ml of water. Use Petri dishes or use scissors to cut down 5 or 6 clear plastic cups to make shallow plastic dishes. Pour enough saltwater to just cover the bottom of each dish (1 for each group). Leave the dishes overnight to evaporate so that new salt crystals will be produced.

Materials for each group

- Black paper
- Salt
- Cup with salt from evaporated saltwater
- Magnifier
- Permanent marker

Materials for each student

- 2 small Styrofoam balls

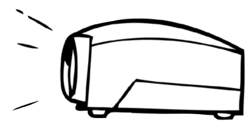
- 2 large Styrofoam balls
- 2 toothpicks

Procedure, Part 1

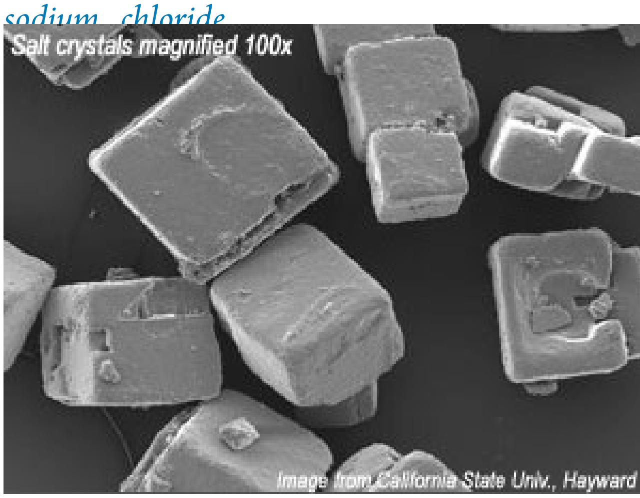
Observe sodium chloride crystals.

1. Place a few grains of salt on a piece of black paper. Use your magnifier to look closely at the salt.
2. Use your magnifier to look at the salt crystals in the cup.

Project the image *Cubic sodium chloride.*

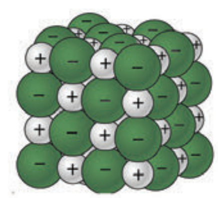


www.middleschoolchemistry.com/multimedia/chapter4/lesson5#cubic_sodium_chloride



inary table salt and a

ake up a salt crystal



Project the animation *Sodium chloride.*

www.middleschoolchemistry.com/multimedia/chapter4/lesson5#sodium_chloride_crystal

The green spheres represent negatively charged chloride ions and the gray spheres represent positively charged sodium ions.

Ask students:

- **What do the photograph, molecular model, and your observations of real salt crystals tell you about the structure of salt?**
  In each case, the salt seems to be shaped like a cube.

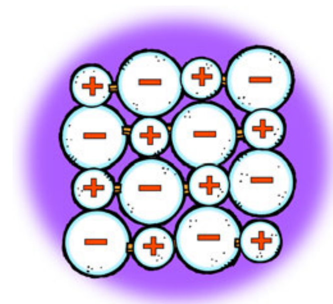
6. Have students build a 3-dimensional model of sodium chloride.

Each student will make 1 unit of sodium chloride. Students in each group will put their sodium chloride units together. You can help the groups combine their structures into a class model of a sodium chloride crystal.

Procedure, Part 2

Make NaCl units.

1. Use the marker to put a "–" on the large balls which represent chloride ions.
2. Use the marker to put a "+" on the small balls, which represent sodium ions.
3. Break two toothpicks in half. Use one of the half-toothpicks to connect the centers of the small and large ions together to make a unit of sodium chloride (NaCl). Do the same thing with the other small and large ball.
4. Use another half-toothpick to connect the two NaCl units in a straight line as shown.

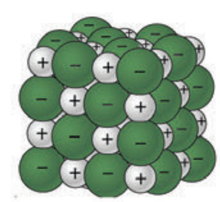


Put NaCl ions together to make one layer of ions.

5. Contribute your line of ions to your group and arrange them to make a 4×4 square of ions.
6. Use half-toothpicks to attach the ends of each line to hold the ions together. You only need to place toothpicks in the balls at the end of each line.

Build a class sodium chloride crystal.

7. Give your group's layer of ions to your teacher. Your teacher will stack these to build a model of a sodium chloride crystal.



Point out that anywhere you look on the crystal, a sodium ion and a chloride ion are always surrounded by the oppositely charged ion. These opposite charges hold the ions together in a crystal.

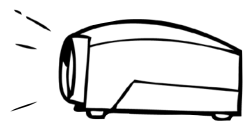
Ask students

- **Based on the way sodium and chloride ions bond together, why are salt crys-**

tals shaped like cubes?
The size and arrangement of the ions forms a cube on the molecular level. Since the pattern repeats over and over again in the same way, the shape stays the same even when the crystal becomes the normal size that we can see.

EXTEND

7. Show students how calcium and chlorine atoms bond to form the ionic compound calcium chloride.



Tell students that there is another common substance called calcium chloride ($CaCl_2$). It is the salt that is used on icy sidewalks and roads. Explain that when calcium and chlorine react they produce ions, like sodium and chlorine, but the calcium ion is different from the sodium ion.

Ask students:

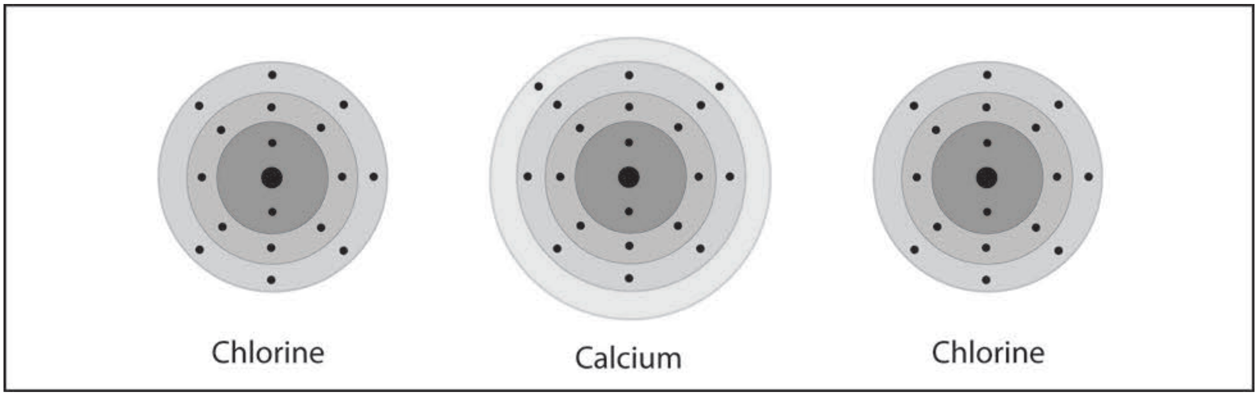
- **What ions do you think $CaCl_2$ is made of?**
  One calcium ion and two chloride ions.

Project the animation *Calcium chloride Ionic Bond*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson5#calcium_chloride_ion
Point out that the calcium loses two electrons, becoming a +2 ion. Each of the two chlorine atoms gains one of these electrons, making them each a –1 ion. Help students realize that 1 calcium ion bonds with 2 chloride ions to form calcium chloride ($CaCl_2$), which is neutral.

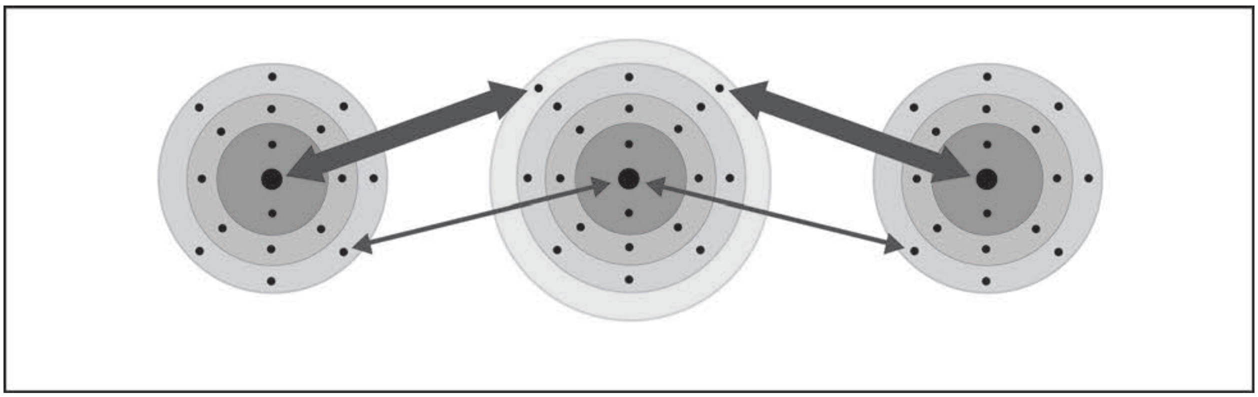
Some atoms gain or lose more than 1 electron. Calcium loses 2 electrons when it becomes an ion. When ions come together to form an ionic bond, they always join in numbers that exactly cancel out the positive and negative charge.

Project the image *Calcium chloride Ionic Bond* .
www.middleschoolchemistry.com/multimedia/chapter4/lesson5#calcium_chloride_ion_2
Review with students the process of ionic bonding covered in the animation.

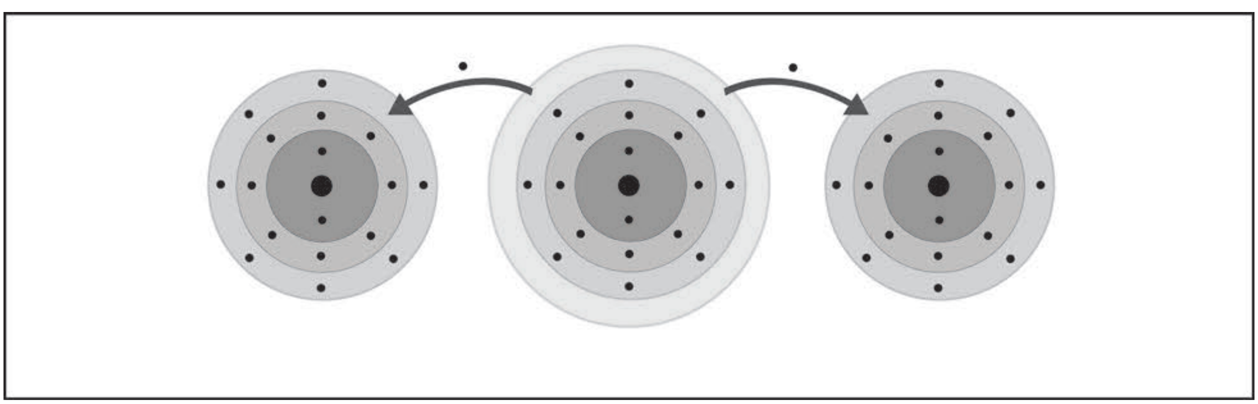
Have students write a short caption beneath each picture to describe the process of ionic bonding in sodium and chloride ions.



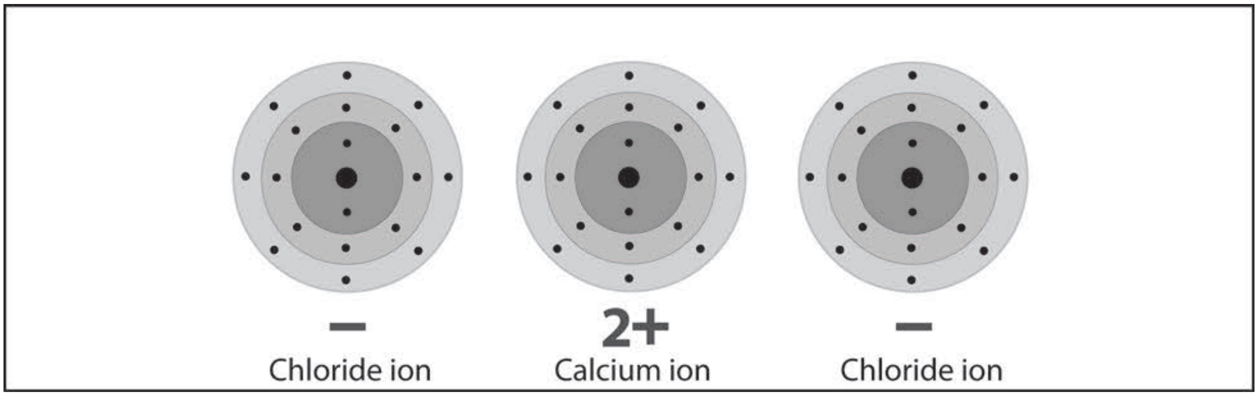
One calcium and two chlorine atoms are near each other.



The protons of the calcium atom attract the electrons from the chlorine atom. The protons of the two chlorine atoms attract the electrons from the calcium atom more strongly as shown by the thicker arrows.

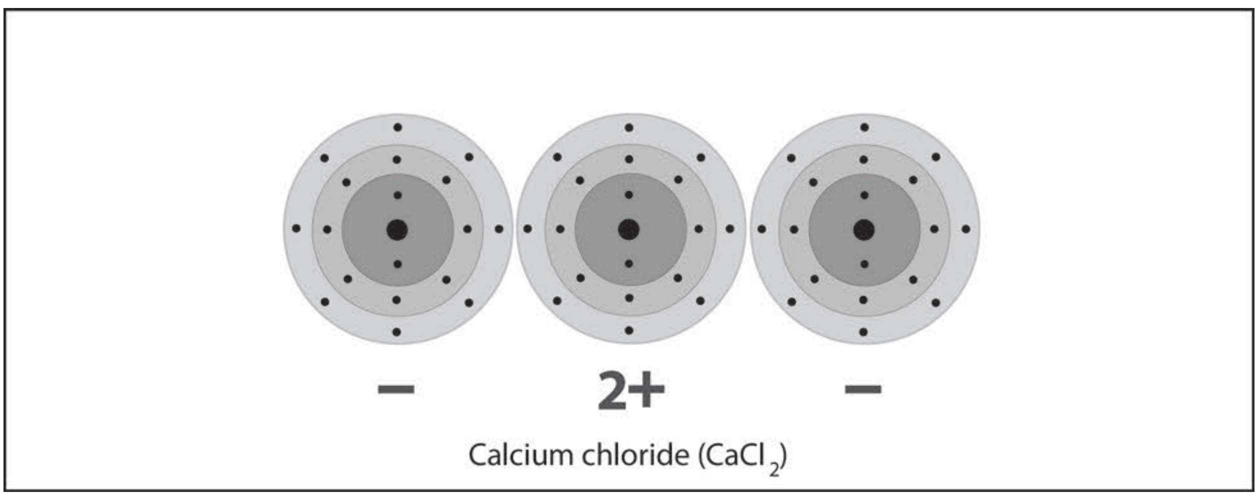


During the interactions between the atoms, the two electrons in calcium's outer energy level are transferred to the outer energy level of each of the chlorine atoms.



Since calcium *lost* two electrons, it has 20 protons, but only 18 electrons. This makes calcium a *positive* ion with a charge of 2+.

Since each chlorine atom *gained* an electron, they each have 17 protons and 18 electrons. This makes each chloride a *negative* ion with a charge of –1.



Oppositely charged ions attract each other, forming an ionic bond. The bonded ions are more stable than the individual atoms were.

Activity Sheet
Chapter 4, Lesson 5
Energy Levels, Electrons, and Ionic Bonding

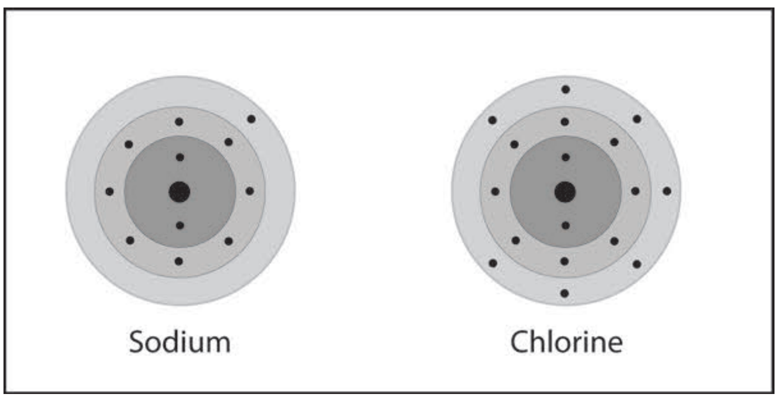
Name ___________________________

Date ___________________________

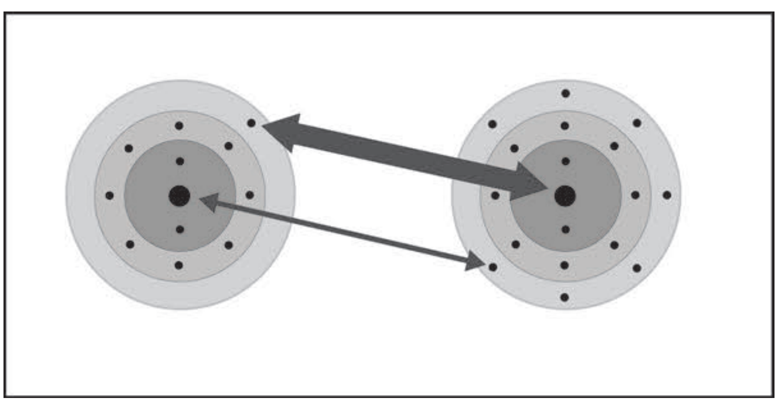
EXPLAIN IT WITH ATOMS & MOLECULES

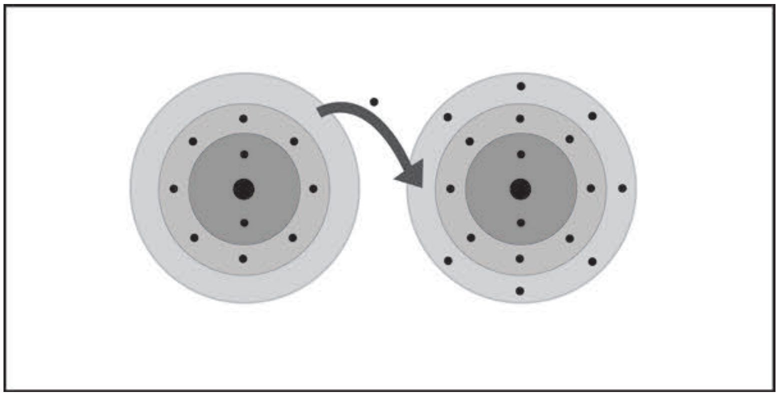
1. What is the basic difference between covalent and ionic bonding?

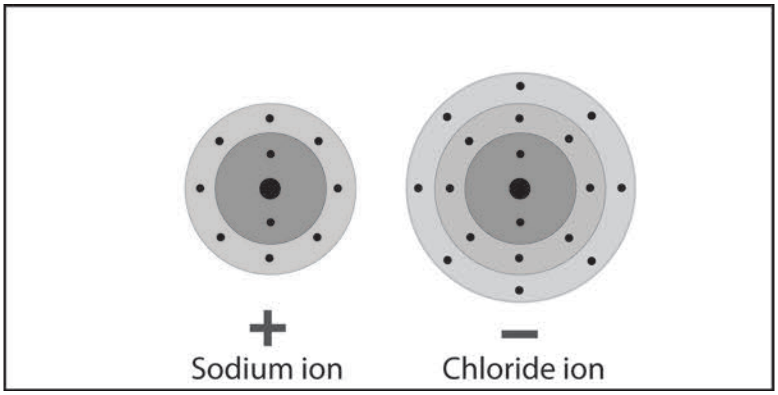
2. Write a short caption beside each picture to describe the process of ionic bonding.

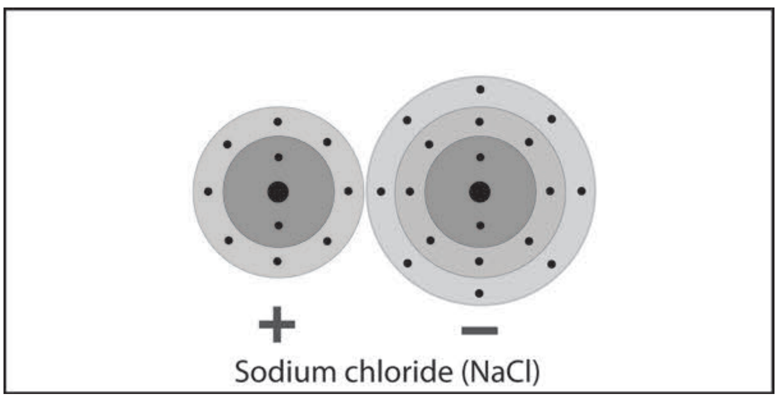


Sodium and chlorine atoms are near each other.









ACTIVITY

Question to investigate
Why are salt crystals cube-shaped?

Materials for each group

- Black paper
- Salt
- Cup with salt from evaporated saltwater
- Magnifier
- Permanent marker

Materials for each student

- 2 small Styrofoam balls
- 2 large Styrofoam balls
- 2 toothpicks

Procedure, Part 1

Observe sodium chloride crystals.

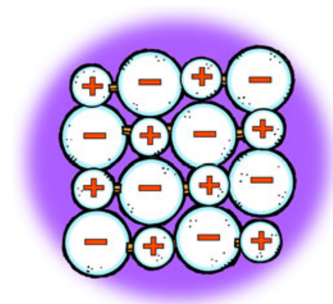
1. Place a few grains of salt on a piece of black paper. Use your magnifier to look closely at the salt.
2. Use your magnifier to look at the salt crystals in the cup.

Procedure, Part 2

Make NaCl units.

1. Use the marker to put a "–" on the large balls, which represent chloride ions.
2. Use the marker to put a "+" on the small balls, which represent sodium ions.
3. Break two toothpicks in half. Use one of the half-toothpicks to connect the centers of the small and large ions together to make a unit of sodium chloride (NaCl). Do the same thing with the other small and large ball.
4. Use another half-toothpick to connect the two NaCl units in a straight line as shown.

Put NaCl ions together to make one layer of ions.



5. Contribute your line of ions to your group and arrange them to make a 4×4 square of ions.
6. Use half-toothpicks to attach the ends of each line to hold the ions together. You only need to place toothpicks in the balls at the end of each line.

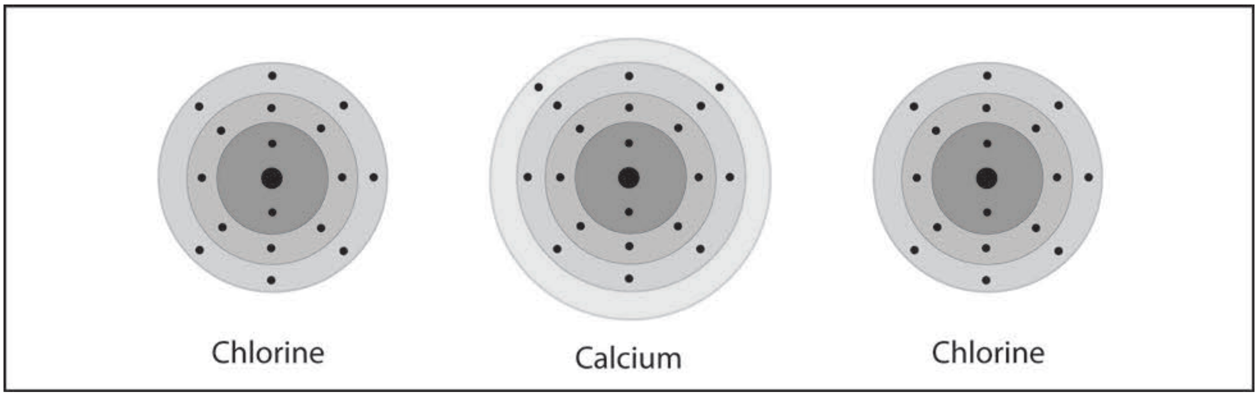
Build a class sodium chloride crystal.

7. Give your group's layer of ions to your teacher. Your teacher will stack these to build a model of a sodium chloride crystal.

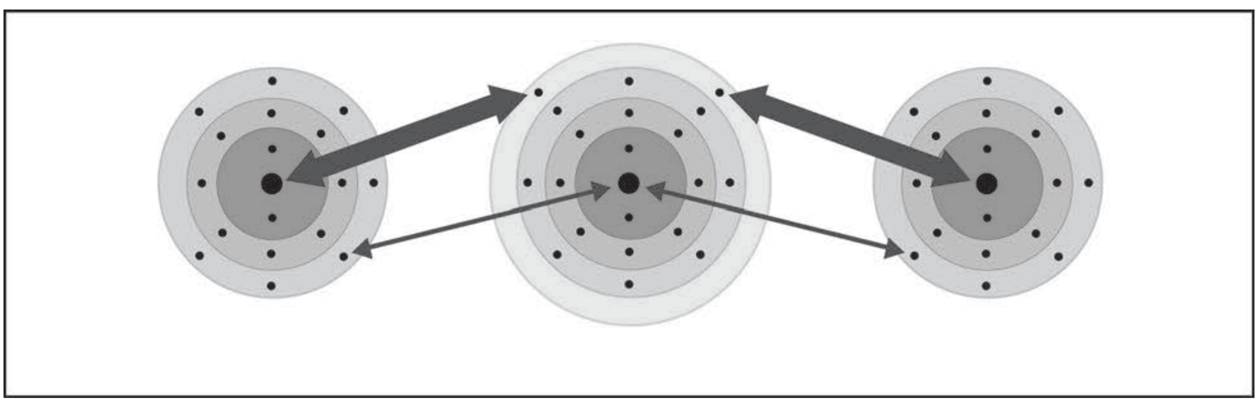
3. Knowing what you do about sodium and chloride ions, why are salt crystals cube-shaped?

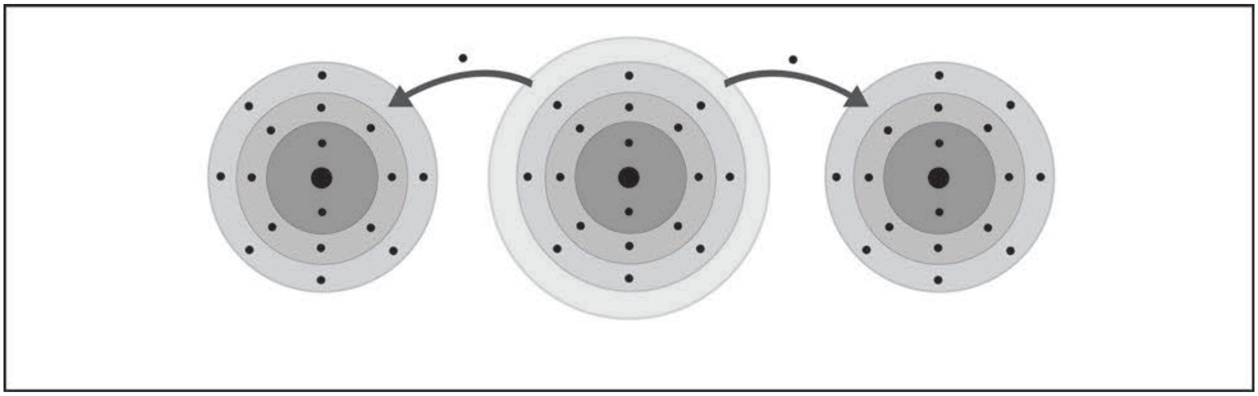
TAKE IT FURTHER

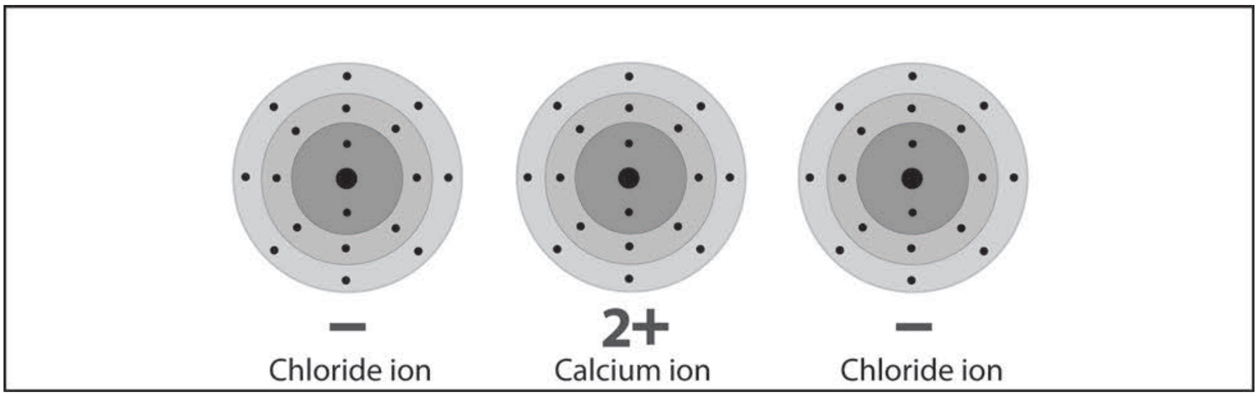
4. Write a short caption beneath each picture below and on the next page to describe the process of ionic bonding. The first one is done for you below.

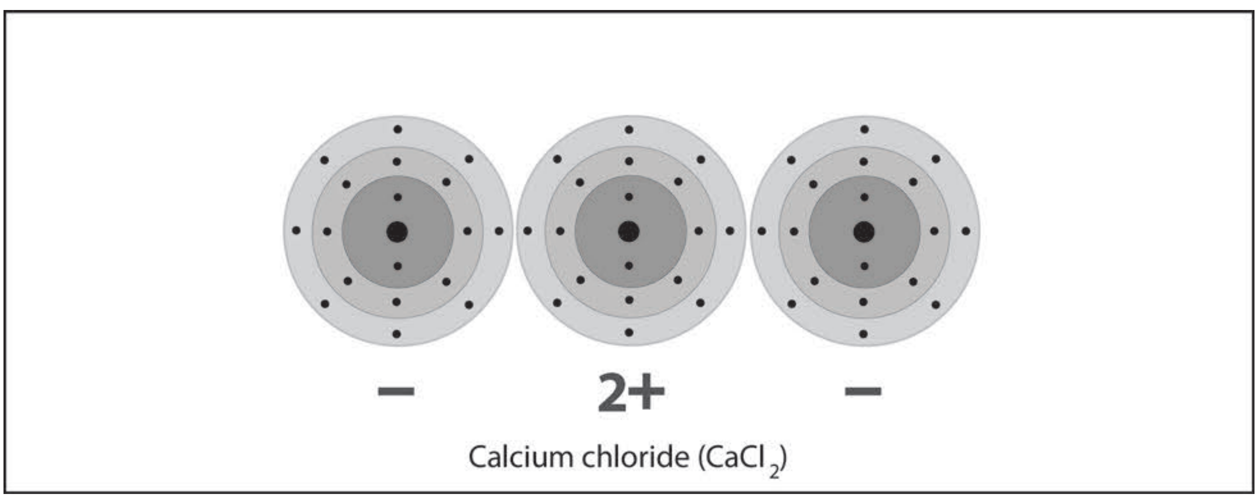


One calcium and two chlorine atoms are near each other.






−
2+
−
Chloride ion
Calcium ion
Chloride ion


−
2+
−
Calcium chloride ($CaCl_2$)

Chapter 4, Lesson 6: Represent Bonding with Lewis Dot Diagrams

If you are required to teach Lewis dot structures, this short lesson can help you extend what students have learned about modeling covalent and ionic bonding. Since there is no hands-on activity component, this lesson is not in a 5-E lesson plan format.

Key Concepts

- There are shorthand ways to represent how atoms form covalent or ionic bonds.
- Lewis dot diagrams use dots arranged around the atomic symbol to represent the electrons in the outermost energy level of an atom.
- Single bonds are represented by a pair of dots or one line between atoms.
- Double bonds are represented by two pairs of dots or two lines between atoms.
- Triple bonds are represented by three pairs of dots or three lines between atoms.

Summary

Students will be introduced to the basics of Lewis dot diagrams as they compare the energy level models used in chapter 4 to dot diagrams. Along with the teacher, they will review the Lewis dot diagrams for a few common covalent and ionic compounds.

Objective

Students will be able to interpret and draw Lewis dot diagrams for individual atoms and both covalent and ionic compounds.

Evaluation

The activity sheet serves as a formative assessment and gives students practice interpreting Lewis dot diagrams. A more formal summative assessment is included at the end of each chapter.

About this Lesson

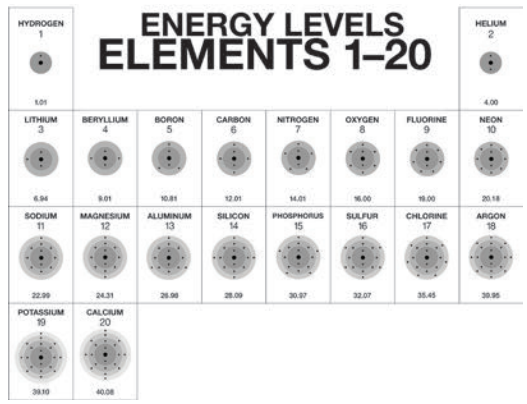
The model of the atom and of covalent and ionic bonding that students have used so far emphasizes the attractions between bonding atoms. The nucleus, electrons, and double-headed arrows show that the protons and electrons from one atom attract the oppositely charged electrons and protons of the other atom, resulting in bonding. The energy levels show that only valence electrons are involved in bonding.

After students understand the important role of attraction of opposite charges, you may introduce them to a common, more symbolic, short-hand way of showing how atoms are bonded together. This information is offered if you feel that showing students these other models of bonding would be useful or if you need to cover basic information about Lewis dot diagrams to satisfy your curriculum.

1. Introduce students to Lewis dot structures.

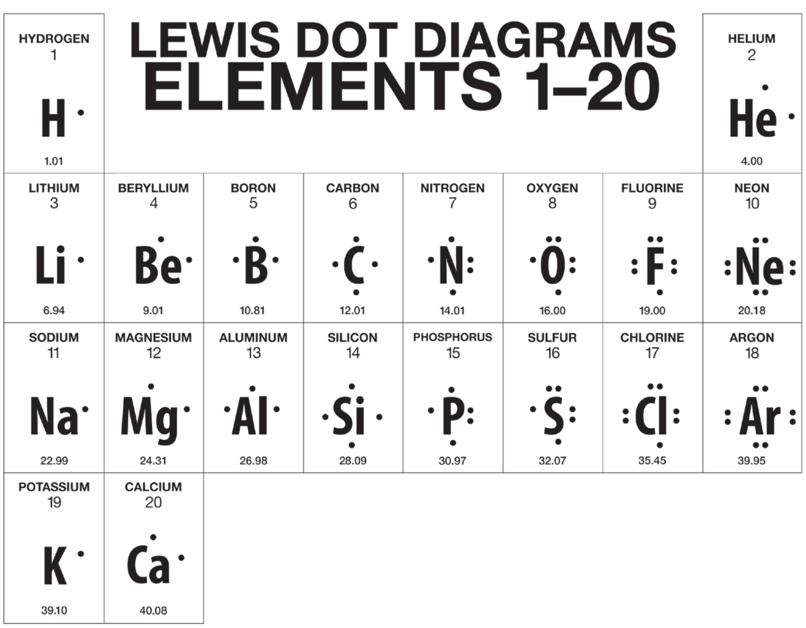
Tell students that one popular method of representing atoms is through Lewis dot diagrams. In a dot diagram, only the symbol for the element and the electrons in its outermost energy level (valence electrons) are shown.

***Note**: In the energy level diagrams students have been using, the electrons are spread out evenly in each energy level. Some books show them spread out this way and some show them in pairs. For Lewis dot structures, they are always shown in pairs. This is to indicate that electrons are in separate orbitals within each energy level. It is not necessary for middle school students to learn about electron orbitals. This information is offered so that it is clearer to you why electrons are often shown in pairs in energy level diagrams and in dot diagrams. An orbital is a 3-dimensional space within an energy level where there is a high probability of finding electrons. The further the energy level is from the nucleus, the more orbitals it has. There can be a maximum of two electrons in each orbital. This is why the electrons are shown in pairs.*



Have students look at the activity sheet for chapter 4, lesson 3 or distribue the energy level chart at the end of this lesson. They will need to compare the energy levels for elements 1–20 that they completed with the chart you show them.

Project the image *Lewis dot diagrams for elements 1–20.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson6#lewis_dot_diagrams

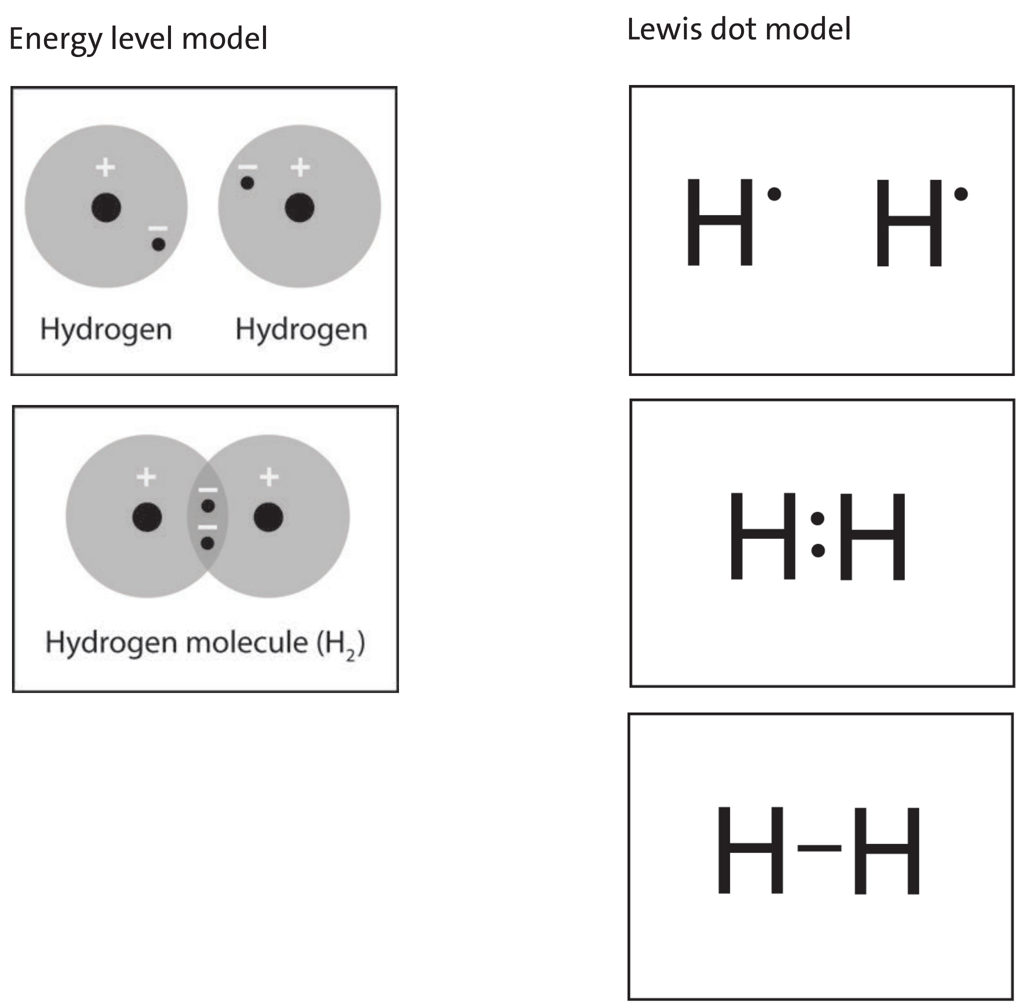


Ask students:

- **Compare the dots around each symbol with the energy levels in your chart. What relationship do you notice between the dots in these two charts?**
  The dots represent the electrons in the outer energy level (valence electrons) from the energy level models.
- **The number of dots near hydrogen and helium are the same as in the energy level chart. Why?**
  The only electrons hydrogen and helium have are valence electrons. All the other dot structures should have fewer electrons than the energy level model because the dot model only shows the outermost electrons.

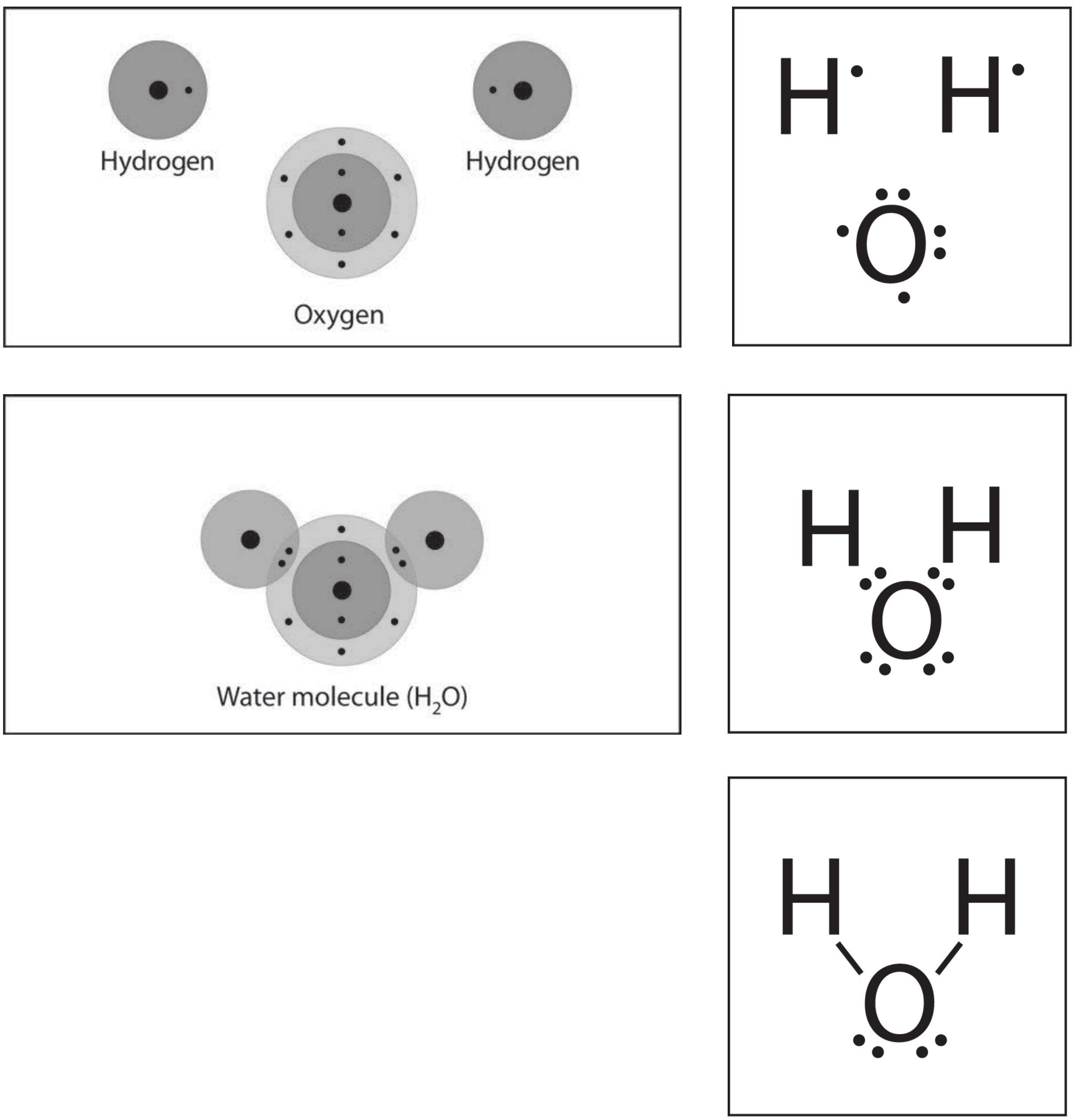
2. Help students recognize the similarities between energy level models and Lewis dot structures that show bonding.

Project the image *Covalent bonding in hydrogen.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson6#covalent_bonding_hydrogen
This image shows both the energy level model and Lewis dot structure of two hydrogen atoms before and after bonding.



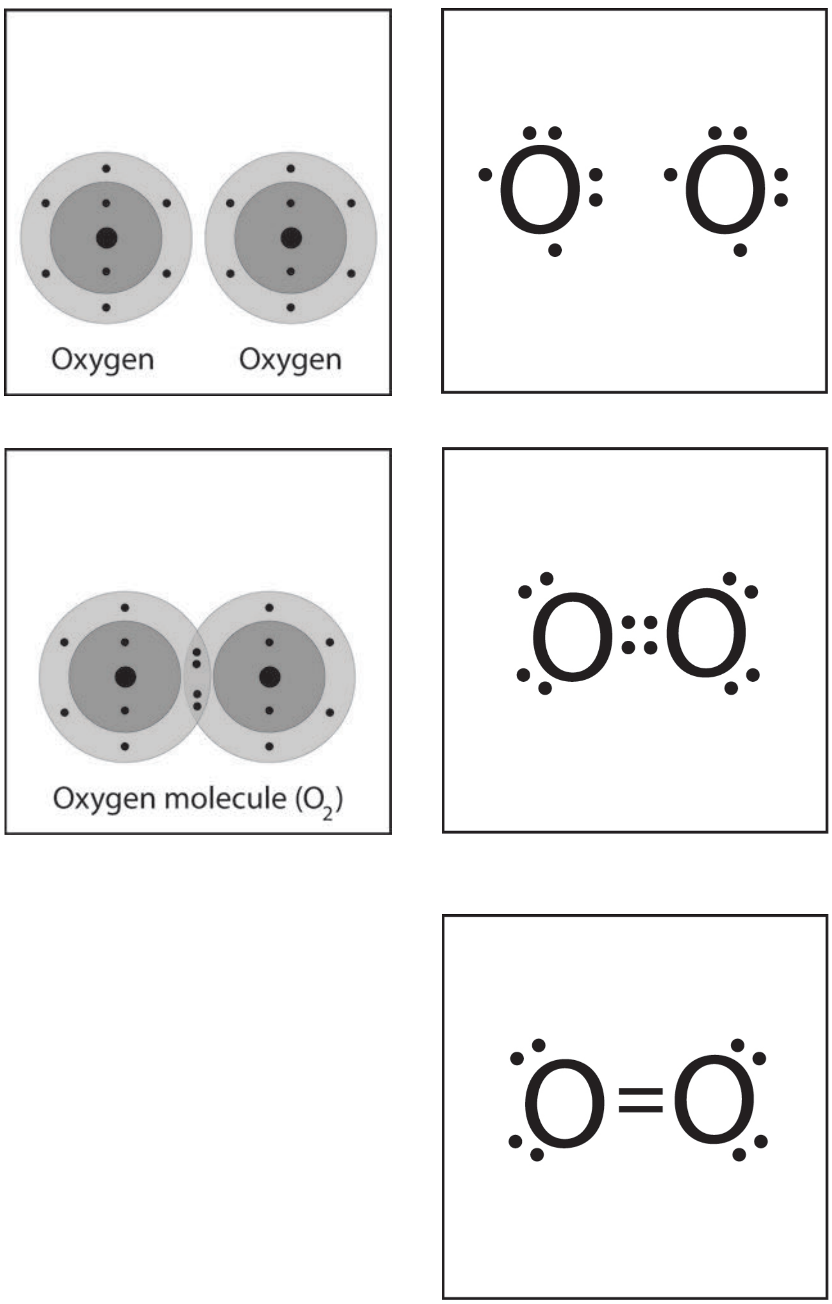
Explain to students that in a dot diagram, the electrons that are shared in the bond are placed between the symbol for each atom. Remind students that the electrons between the two atoms are shared and are counted as belonging to each atom. Show students that in the energy level model for the hydrogen molecule, two electrons are shared. The Lewis dot diagram for the hydrogen molecule also shows that two electrons are shared. There is an even more shorthand approach that shows the bond as a line. The line represents one pair of electrons.

Project the image *Covalent bonding in water.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson6#covalent_bonding_water



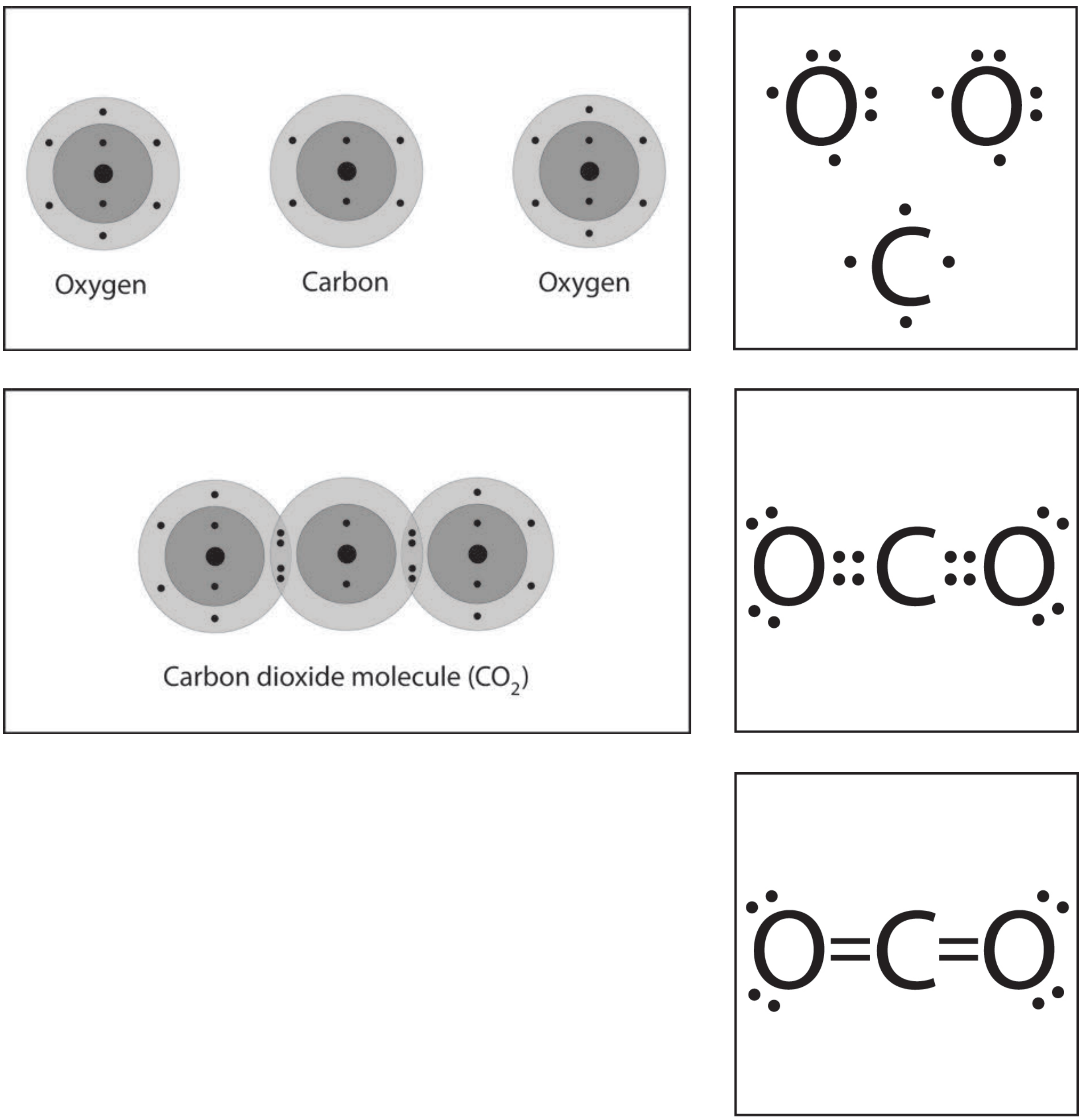
Be sure students notice that the number of dots around the oxygen atom in the Lewis diagram is the same as the number of electrons in the outer energy level of the energy level model. Remind students that the electrons between the atoms are shared and are counted as if they belong to each atom. Show students that in the energy level model for the water molecule, two pairs of electrons are shared. The Lewis dot diagram for the water molecule also shows that two pairs of electrons are shared. The line represents one pair of shared electrons.

Project the image *Covalent bonding in oxygen.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson6/#covalent_bond_oxygen



Show students that in the energy level model for the oxygen molecule, two pairs of electrons are shared. The Lewis dot diagram for the water molecule also shows that two pairs of electrons are shared. The remaining electrons are shown paired up around each oxygen atom. In the alternate Lewis dot diagram, there are two lines because there are two pairs of electrons that are shared.

Project the image *Covalent bonding in carbon dioxide.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson6#covlalent_bonding_carbon_dioxide

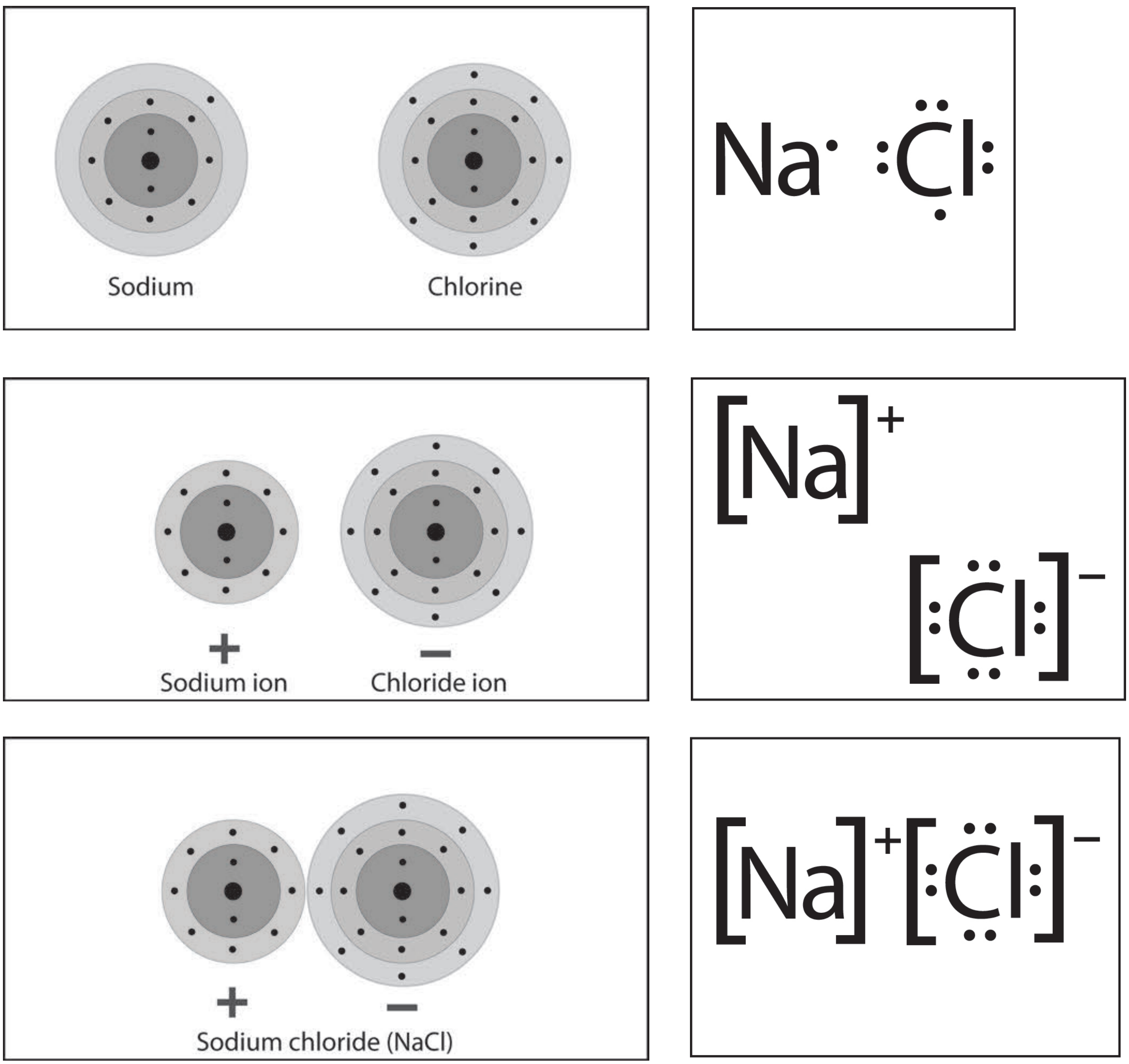


Show students that in the energy level model for carbon dioxide, two pairs of electrons are shared with each oxygen atom. The Lewis dot diagram for carbon dioxide also shows that two pairs of electrons are shared. The remaining electrons are shown paired up around each oxygen atom. In the alternate Lewis dot diagram, there are two lines between each atom to show that two pairs of electrons are shared.

3. Show how Lewis dot diagrams also represent ionic bonding.

Tell students that dot diagrams can also be used to show ionic bonding.

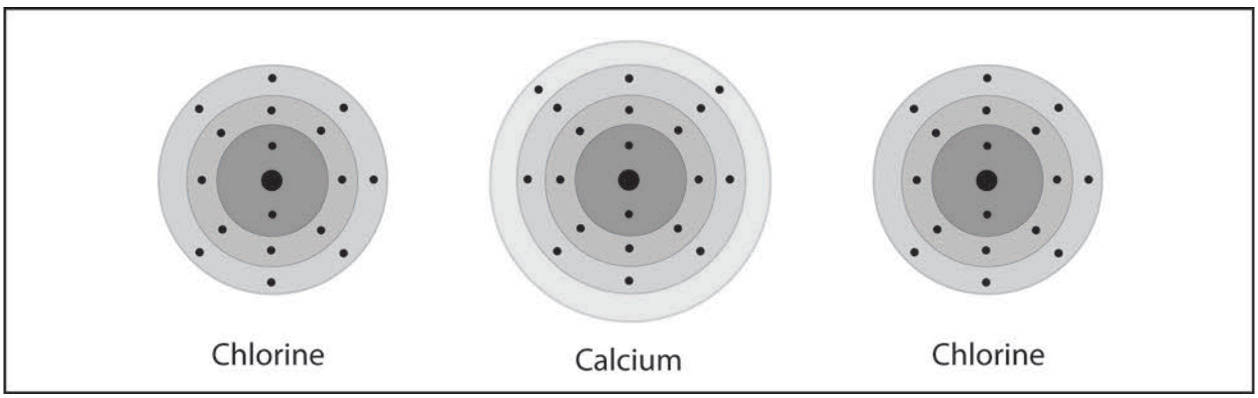
Project the image *Ionic bonding of sodium chloride.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson6#ionic_bond_sodium_chloride

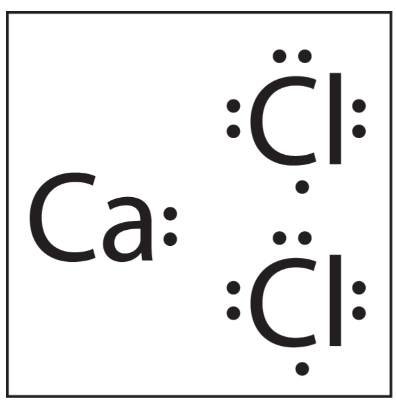


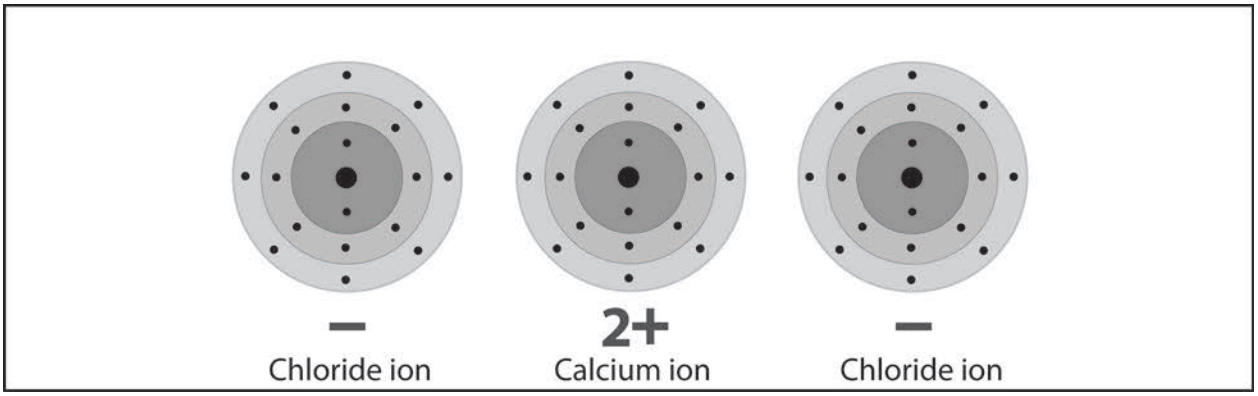
Ask students:

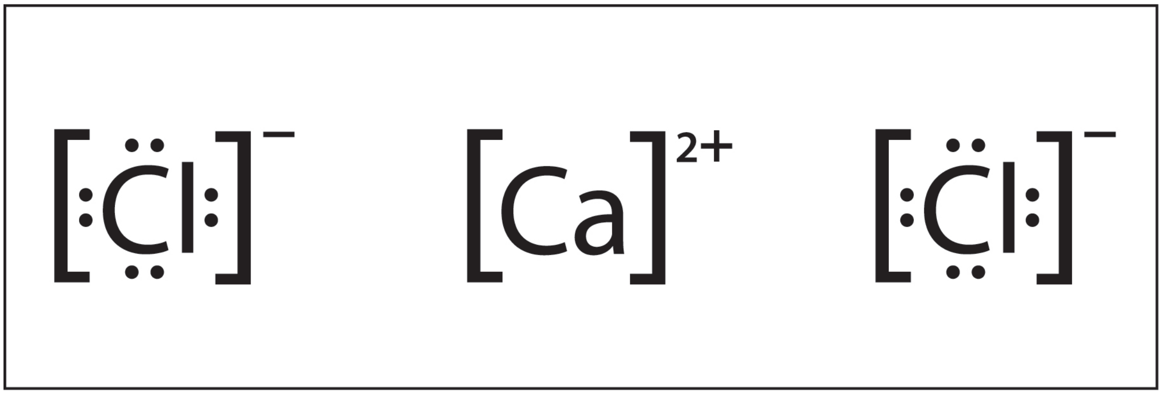
- **In the second dot diagram, why are there no electrons surrounding sodium?**
  The electron was transferred to chlorine. The dot diagram only shows electrons in the atom's outermost energy level. No electrons are shown since the only electron in the outermost energy level of sodium was transferred to chlorine.
- **In the final dot diagram of NaCl, the dots between the sodium and chlorine are between the atoms. Are these atoms sharing the electrons?**
  No. All electrons shown in the dot diagram belong to chlorine. The pair of electrons is between the two letters only because the symbols are shown close to each other to represent the attraction between the oppositely charged ions.

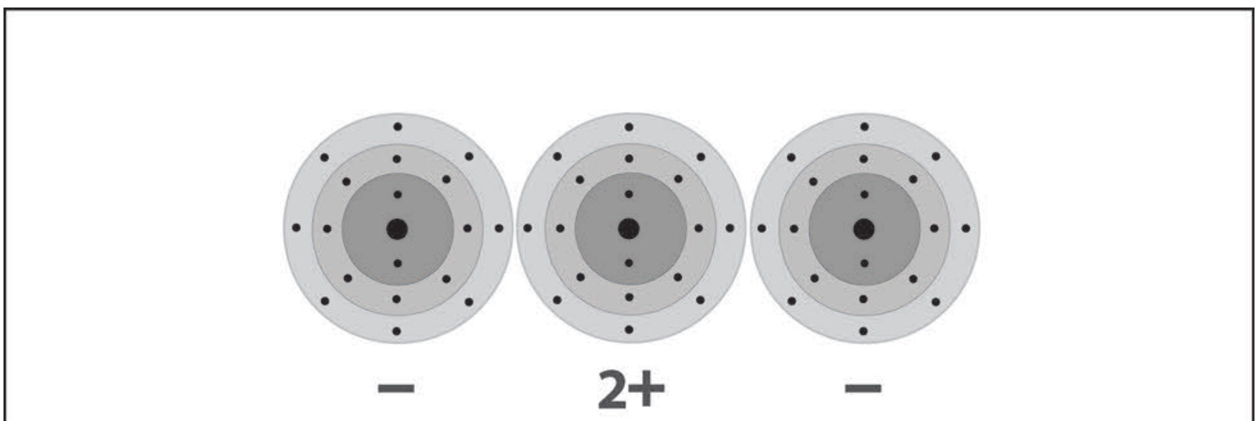
Project the image *Ionic bonding of calcium chloride.*
www.middleschoolchemistry.com/multimedia/chapter4/lesson6#ionic_bonding_calcium_chloride



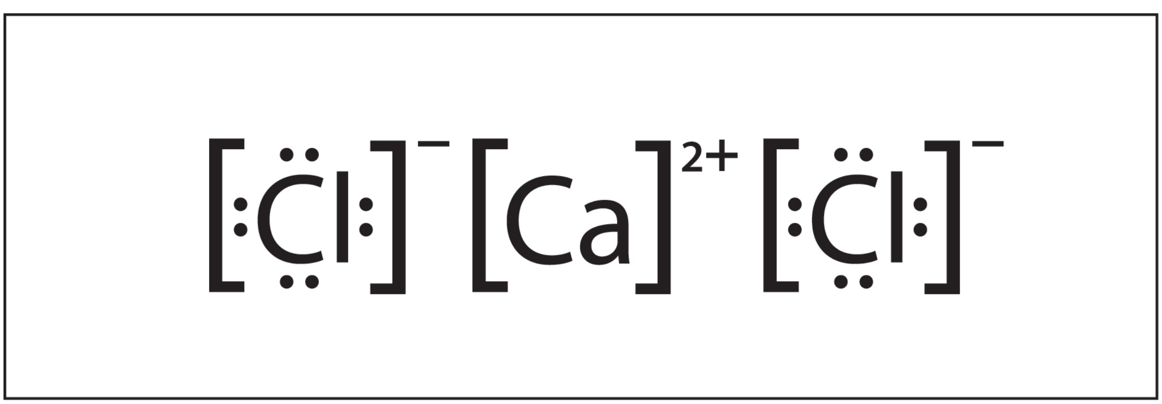








Calcium chloride ($CaCl_2$)



Ask students:

- **What do you think the 2+ near the calcium means?**
  Calcium had two electrons in its outer energy level. It gave each chlorine one of these electons. Because the calcium has two more protons than electrons it has a charge of 2+.
- **In the final dot diagram of $CaCl_2$, the dots between the calcium and chlorine are between the atoms. Are these atoms sharing the electrons?**
  No. All electrons shown in the dot diagram belong to chlorine. The pair of electrons is between the two letters only because the symbols are shown close to each other to represent the attraction between the oppositely charged ions.

Activity Sheet
Chapter 4, Lesson 6
Represent Bonding with Lewis Dot Diagrams

Name ______________________

Date ______________________

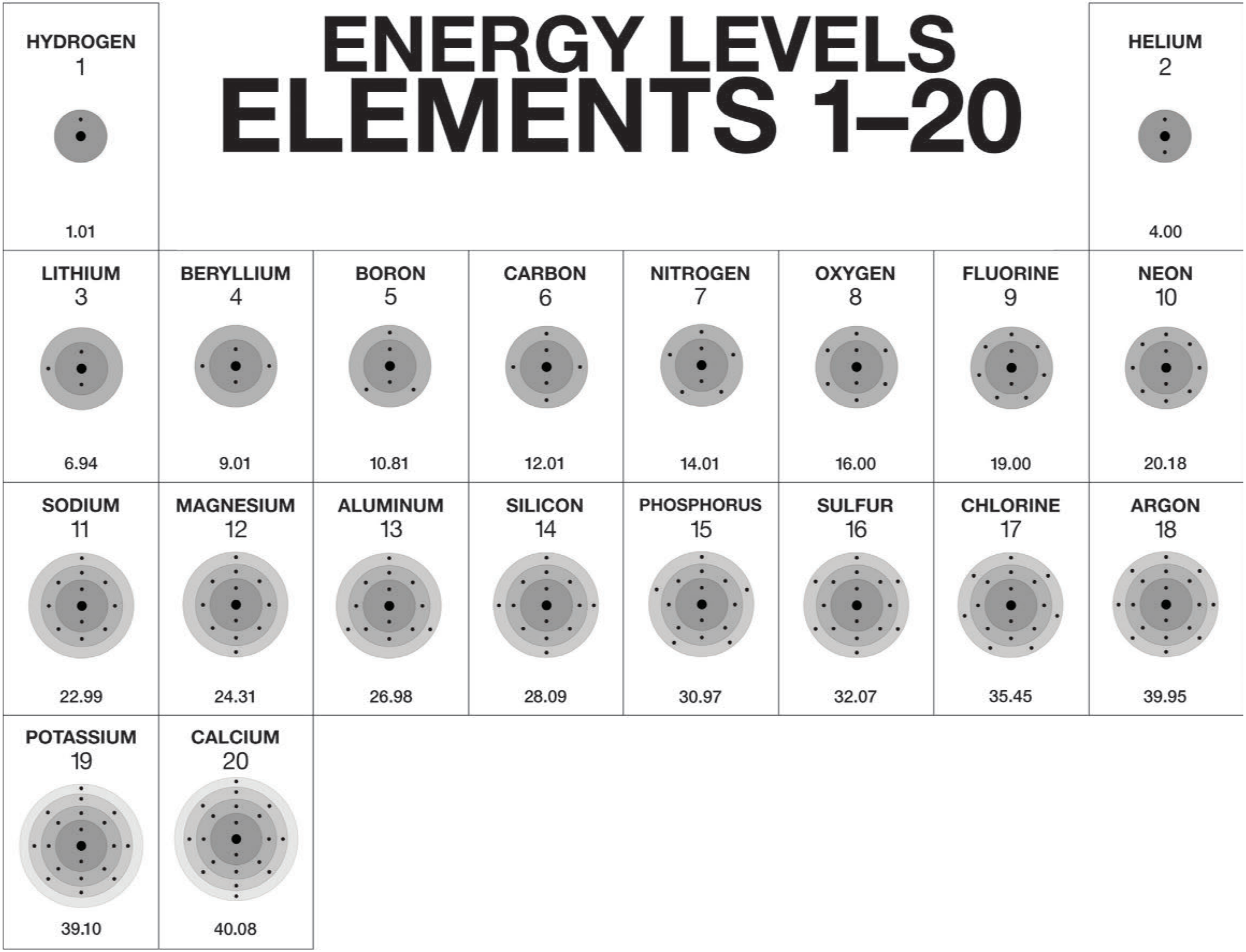
In chapter 4, you saw energy level models for each atom that used concentric circles to represent energy levels and dots for electrons. These diagrams were also used to show what happens to the electrons when different atoms bond. Sometimes electrons were shared (covalent bonding) and sometimes electrons were transferred from one atom to another (ionic bonding).

There is a common, shorthand way to represent bonding called Lewis dot diagrams. Dots still represent electrons, but they are drawn around the symbol for the element. And only the electrons in the outermost energy level are drawn.

1. Compare the periodic table of energy levels to the Lewis dot diagrams. Look at the dots around each symbol and the energy levels in your chart. What relationship do you notice between the dots in these two charts?

2. The number of dots near hydrogen and helium are the same as in the energy level chart. Why?

ENERGY LEVELS ELEMENTS 1–20

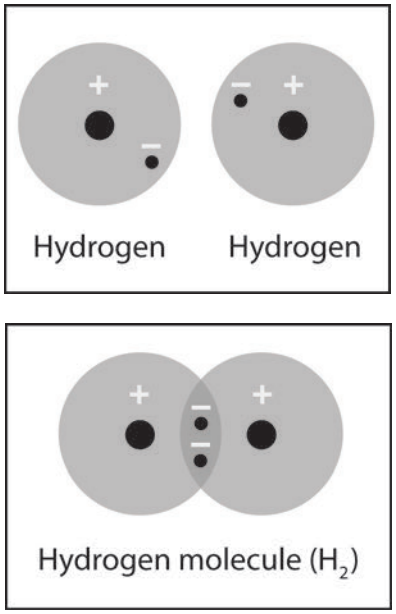


LEWIS DOT DIAGRAMS
ELEMENTS 1–20

HYDROGEN 1 H 1.01							HELIUM 2 He 4.00
LITHIUM 3 Li 6.94	BERYLLIUM 4 Be 9.01	BORON 5 B 10.81	CARBON 6 C 12.01	NITROGEN 7 N 14.01	OXYGEN 8 O 16.00	FLUORINE 9 F 19.00	NEON 10 Ne 20.18
SODIUM 11 Na 22.99	MAGNESIUM 12 Mg 24.31	ALUMINUM 13 Al 26.98	SILICON 14 Si 28.09	PHOSPHORUS 15 P 30.97	SULFUR 16 S 32.07	CHLORINE 17 Cl 35.45	ARGON 18 Ar 39.95
POTASSIUM 19 K 39.10	CALCIUM 20 Ca 40.08						

Covalent bonding in the hydrogen molecule, H_2

Energy level model



Lewis dot model

H· H·

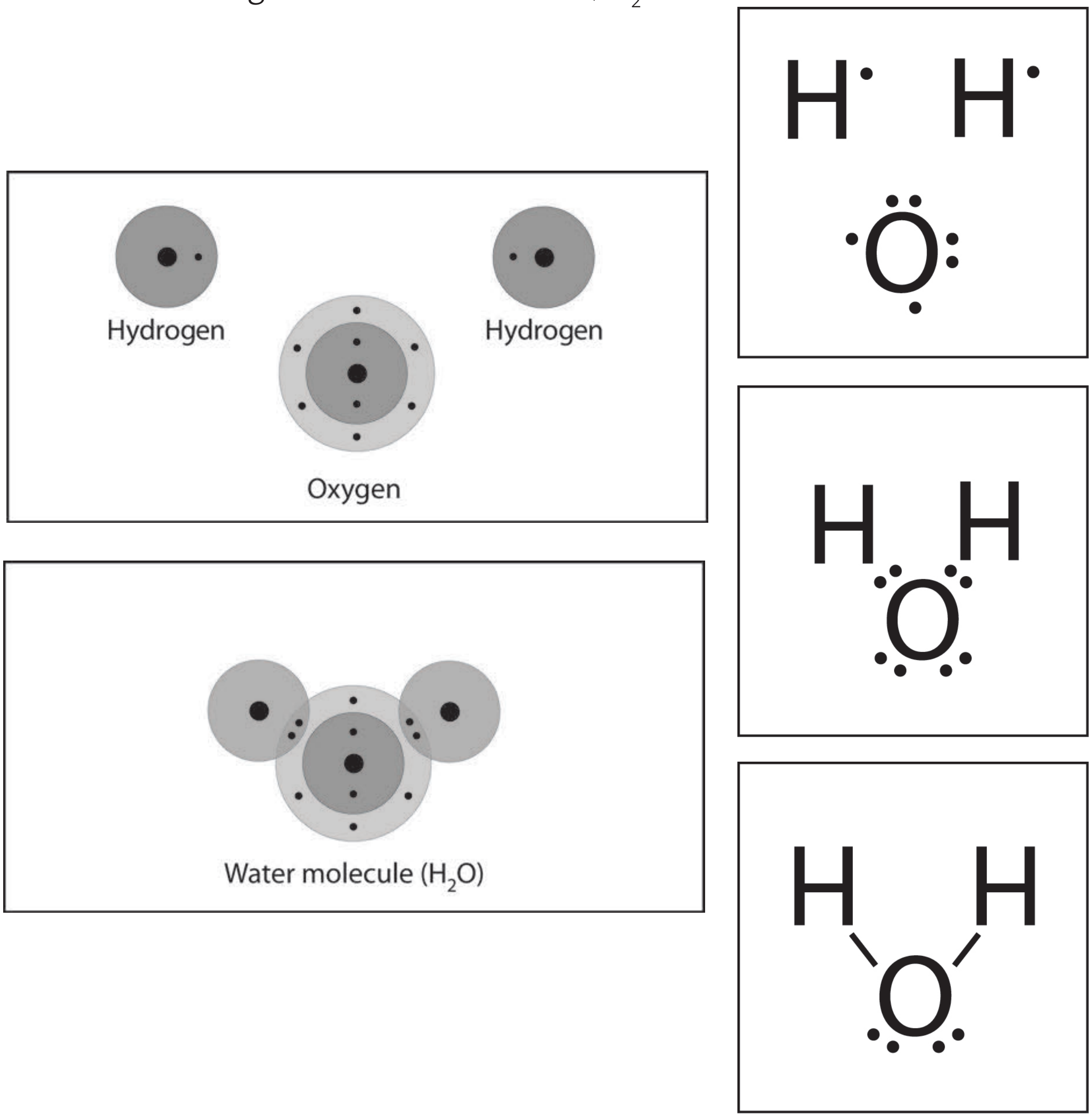
H:H

H–H

3. What do the pair of dots between the two letters "H" represent?

4. What does the line between the two letters "H" represent?

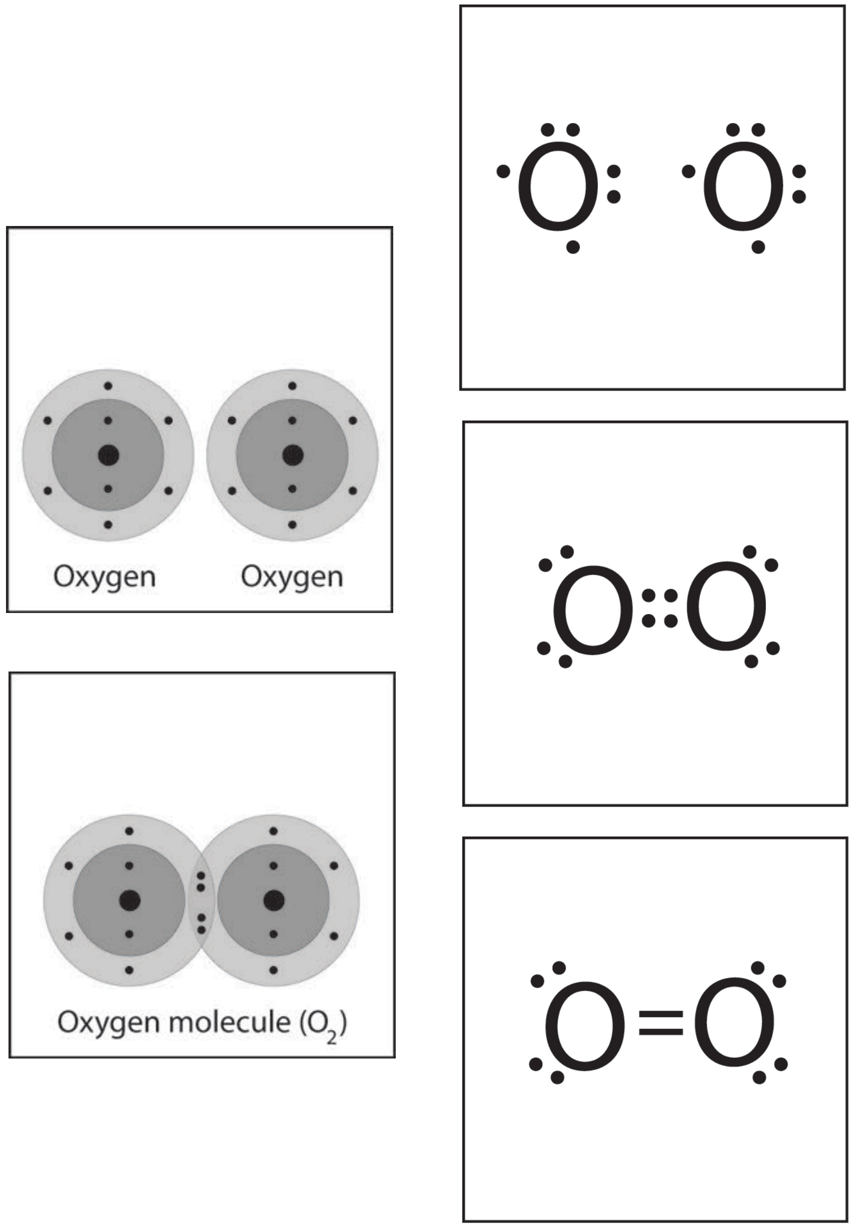
Covalent bonding in the water molecule, H_2O



5. Water has two hydrogen atoms covalently bonded to an oxygen atom. Methane has four hydrogen atoms covalently bonded to a carbon atom.

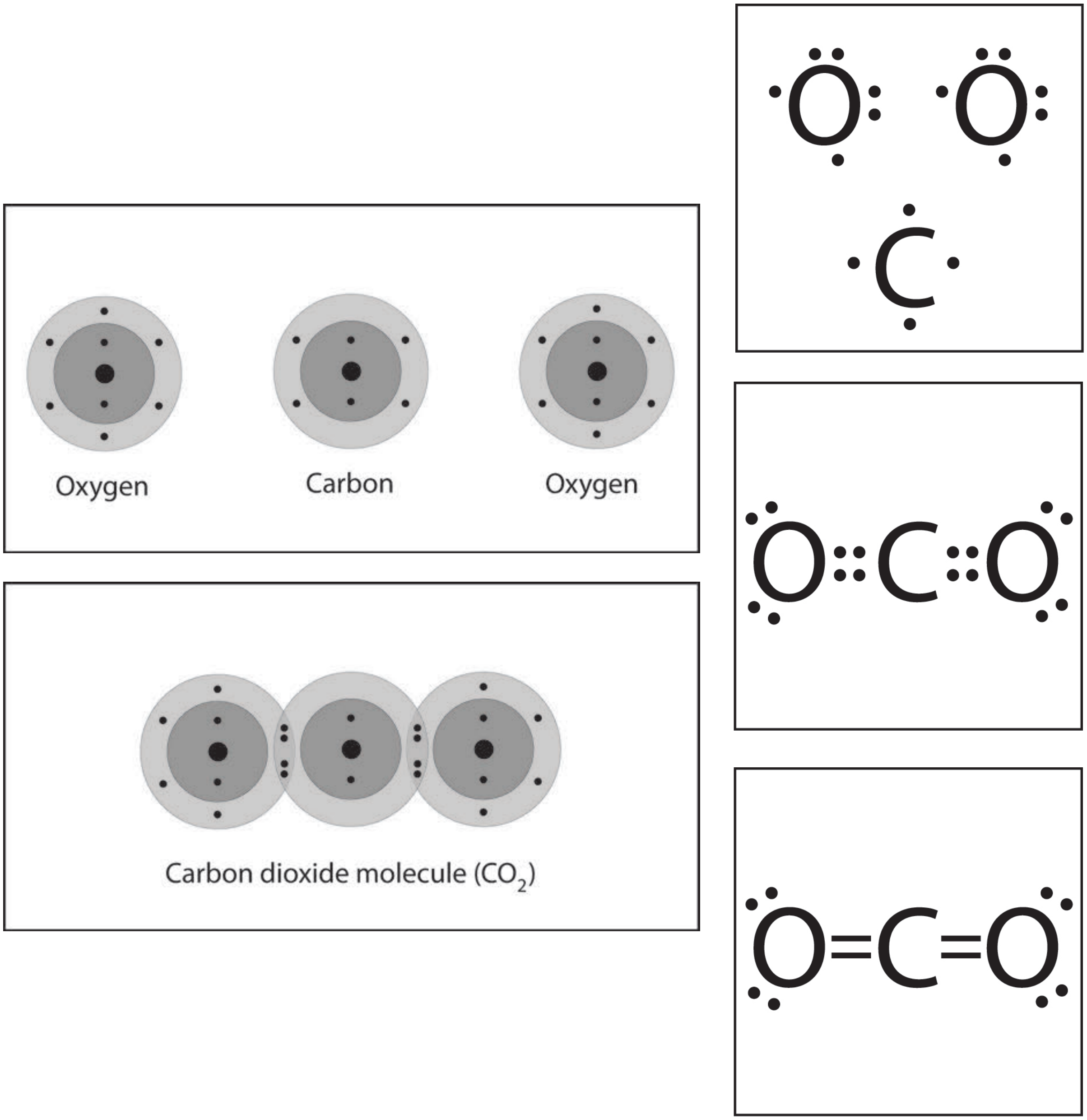
   Draw Lewis diagrams for methane using dots for pairs of electrons and then lines for pairs of electrons.

Covalent bonding in the oxygen molecule, O_2



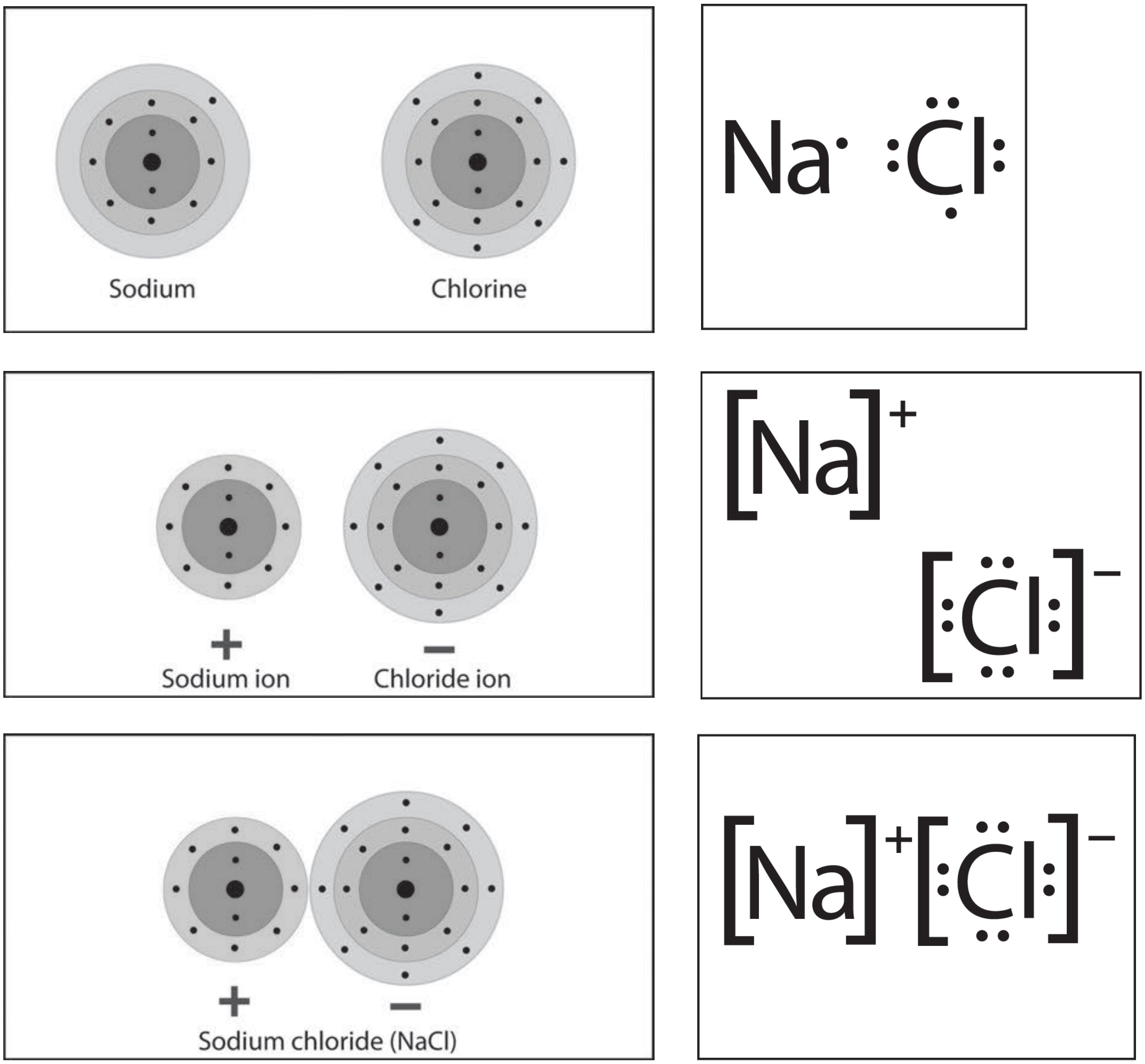
6. Why are there four dots between the two oxygen atoms?

Covalent bonding in the carbon dioxide molecule, CO_2



7. Why are there two sets of lines between the carbon and each oxygen atom?

Ionic bonding of sodium chloride, NaCl

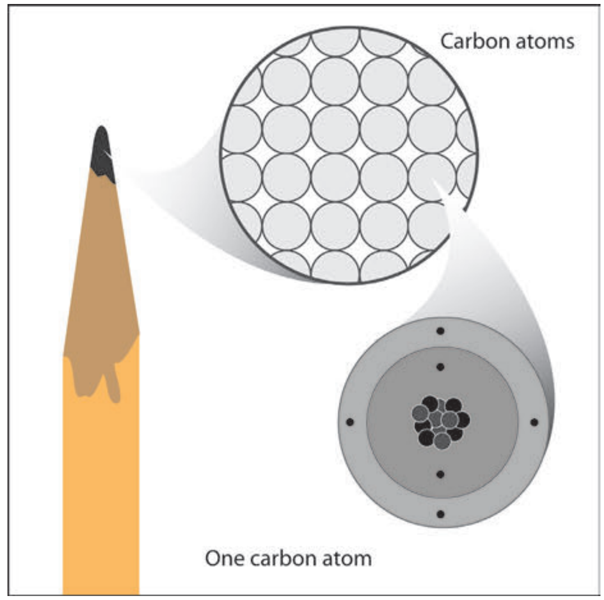


8. In the second dot diagram, why are there no electrons surrounding sodium?

9. In the final dot diagram of NaCl, the dots between the sodium and chlorine are between the atoms. Are these atoms sharing the electrons?

Chapter 4—Student Reading

Parts of the atom

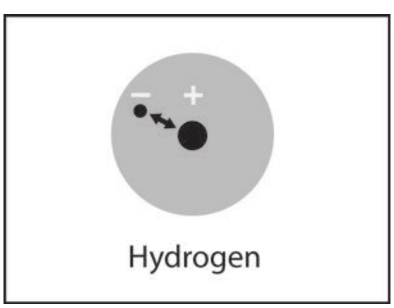


An atom is made up of *protons*, *neutrons*, and *electrons*. Look at the model of a carbon atom from the graphite in the point of a pencil. Protons and neutrons are in the center or *nucleus* of the atom. Electrons are in regions surrounding the nucleus. In the carbon atom, there are six protons, and six electrons. The vast majority of carbon atoms also have six neutrons.

A proton has a positive charge. An electron has a negative charge. A neutron has no charge. The charge on the proton and electron are exactly the same size but opposite. The same number of protons and electrons exactly cancel each other in a neutral atom.

Two protons push each other away or *repel*. Two electrons also repel each other. But a proton and an electron move toward or *attract* each other. Another way of saying this is that the same or "like" charges repel one another and opposite charges attract one another. Since opposite charges attract each other, the negatively charged electrons in an atom are attracted to the positively charged protons. This attraction is what holds an atom together.

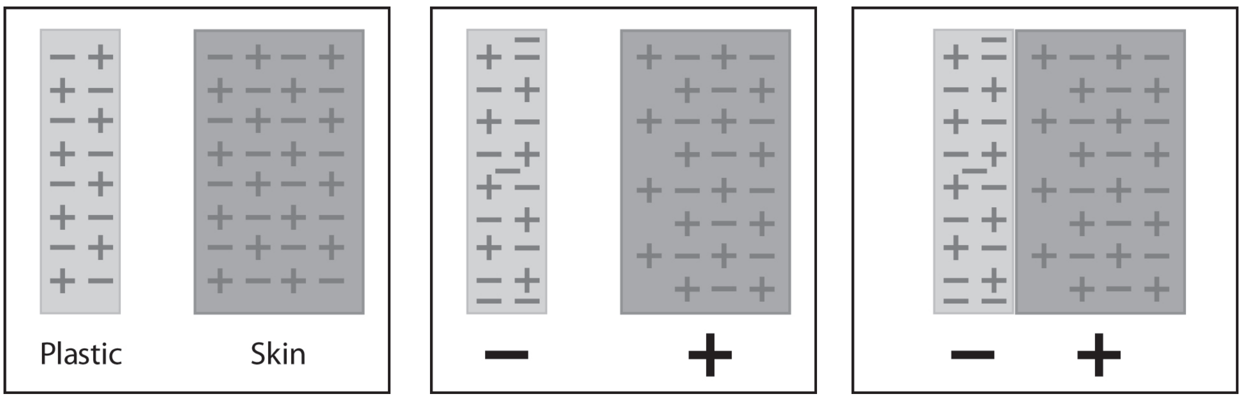
This is a simple model of a hydrogen atom which has one proton and one electron. The arrow shows that the electron is attracted to the proton.



Another model of the hydrogen atom shows a cloudy-looking region in the space surrounding the nucleus. This model represents the electron as a cloud to show that it is not possible to know the exact location of an electron. The electron cloud shows the region surrounding the nucleus where the electron is most likely to be.

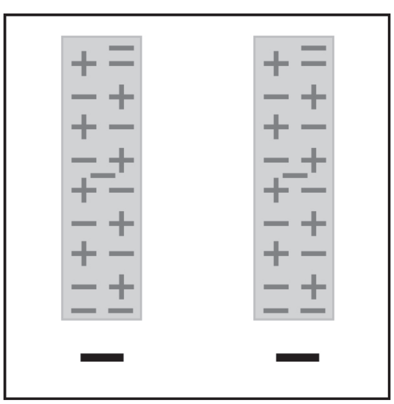
Proton, electrons, and static electricity

You can see evidence of electrons and protons attracting or repelling each other when you make static electricity. For example, when you rub a plastic strip between your fingers, electrons move from your skin to the plastic. If you assume that the plastic and the skin were both neutral before rubbing, the plastic now has more electrons or negative charges than positive.



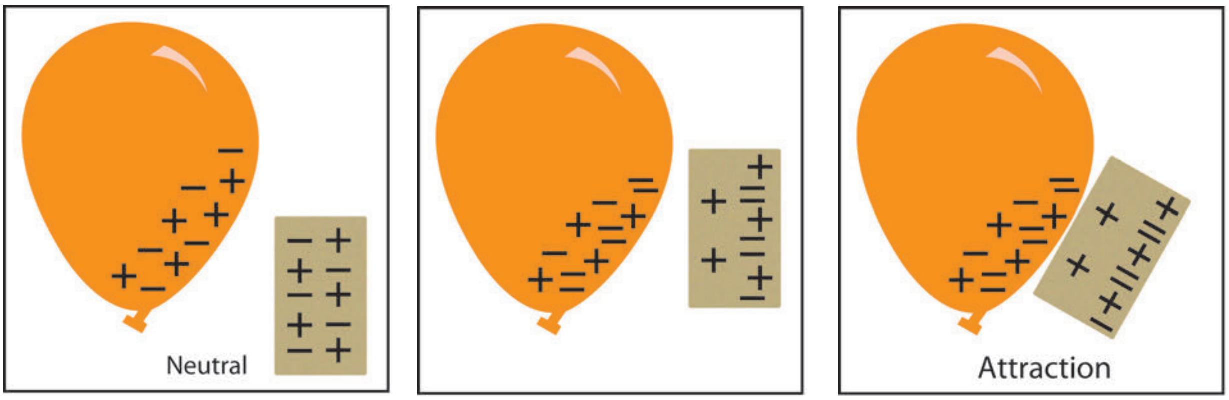
This gives the plastic an overall or *net* negative charge. Since your skin lost some negative charge, it now has more positive charge than negative, so your skin has an overall or net positive charge. When you bring the plastic near your fingers, the plastic is attracted because opposite charges attract.

If you rub *two* plastic strips on your fingers, each strip gains electrons so each one has a net negative charge. If you bring the strips near each other, they repel because like charges repel.



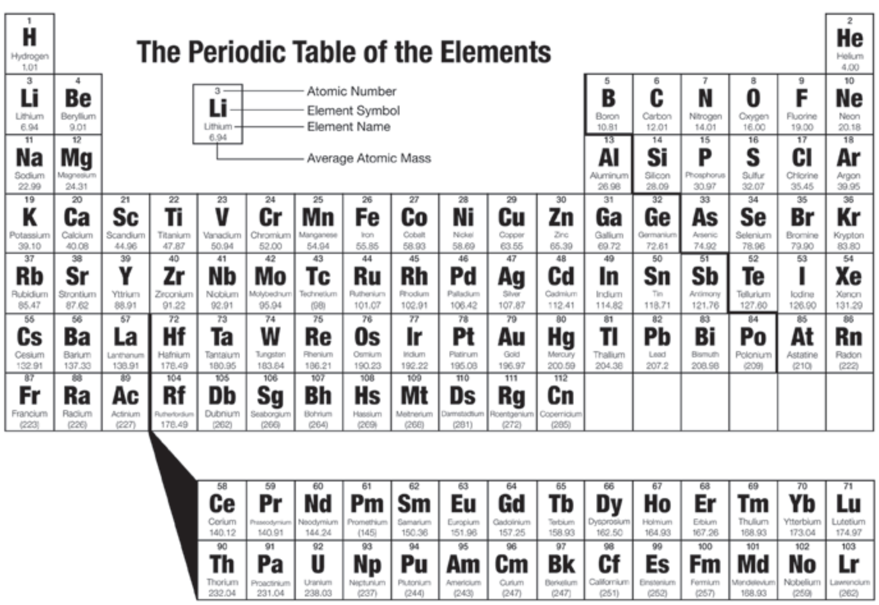
Rubbing a balloon and sticking it to a wall or using it to attract little pieces of paper is also evidence that protons and electrons have opposite charge. When you rub a balloon on your hair or clothes, electrons move onto the balloon. This gives the balloon a negative charge.

When the balloon is brought near a little piece of paper, electrons on the balloon repel electrons in the paper. The electrons in the paper move away from the balloon and leave an area of positive charge near the balloon. The positively charged area of the paper is attracted to the negative balloon and the paper moves to the balloon.

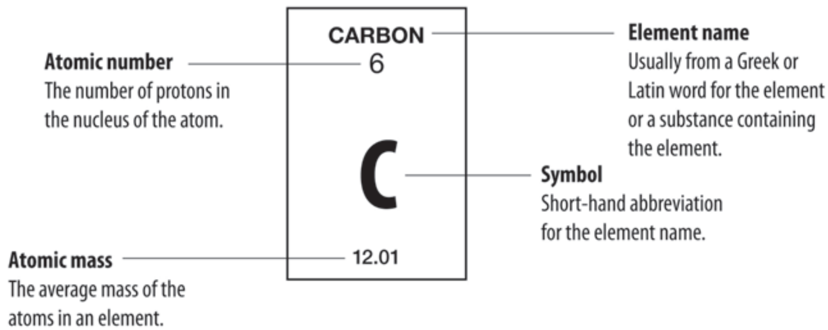


The Periodic Table of the Elements

You have read about protons and electrons, and about the atoms and molecules in different substances. The atoms that make up all solids, liquids, and gases are organized into a chart or table called the periodic table of the elements. The periodic table shows all the atoms that everything in the known universe is made from. Each box contains information about a different atom. It's like the alphabet in which only 26 letters, in different combinations, make up thousands of words. The 100 or so atoms of the periodic table, in different combinations, make up millions of different substances.



Each box in the periodic table contains basic information about an element.



The meaning of the terms "atom" and "element" can be confusing because they are often used as if they are the same thing. They are related to one another but they are not the same. An atom is the smallest particle or "building block" of a substance. An element is a substance made up of all the same type of atom. For instance, a piece of pure carbon is made up of only carbon atoms. The piece of pure carbon is a sample of the element carbon. The people who developed the periodic table could have called it the Periodic Table of the Atoms but they did not have a firm understanding of atoms at that time. Since they were working with actual samples of elements such as copper, mercury, sulfur, etc., they called it the periodic table of the elements.

Atomic mass

The element name, atomic number, and symbol are pretty easy to understand. The atomic mass is a little trickier. The atomic mass of an element is based on the mass of the atoms that make up the element. The mass of the atoms is based on the protons, neutrons, and electrons of the atoms. The mass of the proton and neutron are about the same but the mass of the electron is much smaller (about 1/2000 the mass of the proton or neutron). The vast majority of the atomic mass is contributed by the protons and neutrons.

For any element in the periodic table, the number of electrons in an atom always equals the number of protons in the nucleus. But this is not true for neutrons. Atoms of the same element can have different numbers of neutrons than protons. Atoms of the same element with different numbers of neutrons are called isotopes of that element. The atomic mass given in the periodic table is an *average* of the atomic mass of the isotopes of an element.

For example, the vast majority of carbon atoms have 6 protons and 6 neutrons, but a small percentage of carbon atoms have 6 protons and 7 neutrons, and an even smaller percentage have 6 protons and 8 neutrons. Since the vast majority of carbon atoms have a mass very close to 12, and only a small percentage are greater than 12, the average atomic mass is slightly greater than 12 (12.01). For the atoms of the first 20 elements, the number of neutrons is either equal to or slightly greater than the number of protons.

Hydrogen is an exception to this rule. All hydrogen atoms have one proton but the vast majority have 0 neutrons. There is a small percentage of hydrogen atoms that have 1 neutron and a smaller percentage that have 2 neutrons. When you take the average mass of all the different isotopes of hydrogen, the mass is slightly greater than one (about 1.01).

Electrons are in energy levels surrounding the nucleus

Electrons surround the nucleus of an atom in three dimensions making atoms spherical. Electrons are in different regions around the nucleus like concentric spheres. These regions are called energy levels. Since it is very difficult to draw concentric spheres, the energy levels are usually shown in 2 dimensions.

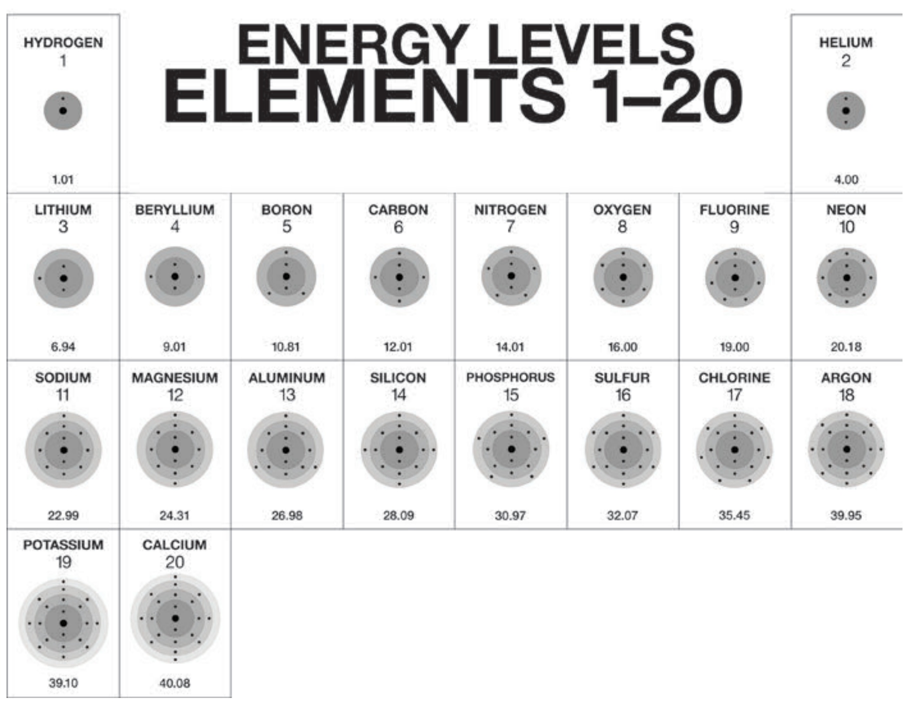
This energy level model represents an oxygen atom. The nucleus is represented by a dot in the center which contains both protons and neutrons. The smaller dots surrounding the nucleus represent electrons in the energy levels. You can tell that this model is oxygen because there are a total of 8 electrons. Since neutral atoms in the periodic table have the same number of electrons as protons, this atom must have 8 protons. The number of protons is the same as the atomic number, so this atom's atomic number is 8, which is oxygen.

Arrangement of elements in the periodic table

There is a limit to the number of electrons that can go into the different energy levels of an atom. A certain number of electrons go into an energy level before they begin to go into the next level. After the first energy level contains 2 electrons (helium), the next electrons go into the second energy level. After the second energy level has 8 electrons (neon), the next electrons go into the third energy level.

After the third energy level has 8 electrons (argon), the next 2 electrons go into the fourth energy level. An energy level model is shown in the chart below for the first twenty elements in the periodic table.

The rows going across the periodic table are called *periods*. The columns going up and down are called *groups* or *families*.



Number of energy levels in each period

- The atoms in the first period have electrons in 1 energy level.
- The atoms in the second period have electrons in 2 energy levels.
- The atoms in the third period have electrons in 3 energy levels.
- The atoms in the fourth period have electrons in 4 energy levels.

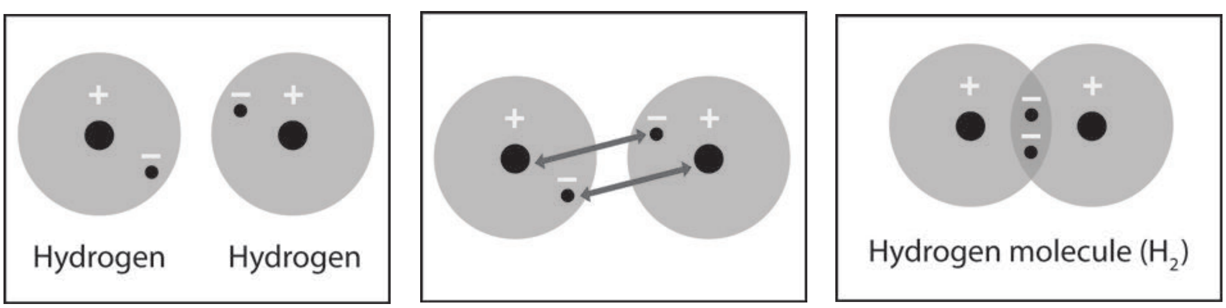
Atoms in a group have the same number of valence electrons

If you look at the atoms in a group, you will see that they each have the same number of electrons in their outermost energy level. Electrons in this level are called valence electrons. For instance, hydrogen, lithium, sodium, and potassium all have 1 valence electron. Valence electrons are important because they interact with other atoms and are responsible for many of the characteristic properties of the atom.

HOW ATOMS BOND TO EACH OTHER

Covalent bonding

Remember that a hydrogen atom has 1 proton and 1 electron and that the electron and the proton are attracted to each other. But if the atoms get close enough to each other, the electron from each hydrogen atom feels the attraction from the proton of the other hydrogen atom (shown by the double headed arrow).



The attractions are not strong enough to pull the electron completely away from its own proton. But the attractions are strong enough to pull the two atoms close enough together so that the electrons feel the attraction from both protons. When the electrons are attracted to and shared by both atoms, the individual hydrogen atoms have bonded to become the *molecule* H_2. This type of bond is called a *covalent* bond. In a covalent bond, electrons from each atom are attracted or "shared" by *both* atoms. Two or more atoms covalently bonded are called a *molecule*.

There are two main requirements for atoms to form a covalent bond and make a molecule:

- There needs to be a strong enough attraction between the electrons in each atom for the protons in the other atom.
- There needs to be room in the outer energy level of both atoms.

Once bonded, the hydrogen molecule is more stable than the individual hydrogen atoms. By being part of a covalent bond, the electron from each hydrogen atom gets to be near two protons instead of only the one proton it started with. Since the electrons are closer to more protons, the molecule of two bonded hydrogen atoms is more stable than the two individual unbonded hydrogen atoms.

Atoms bond until their outer energy levels are full

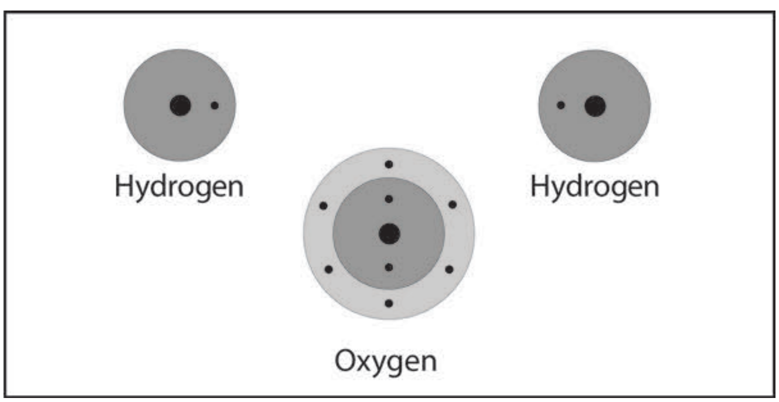
The two electrons in the hydrogen molecule (H_2) can be thought of as "belonging" to each atom. This means that each hydrogen atom now has two electrons in its first energy level. The first energy level is the outer energy level for hydrogen and can only accommodate or "hold" two electrons. This means that the outer energy level is full. Atoms will covalently bond to one another until each atom's outer energy level is full.

Once the outer energy levels are full, additional atoms will not covalently bond to the atoms in the H_2 molecule. This will not happen for two main reasons:

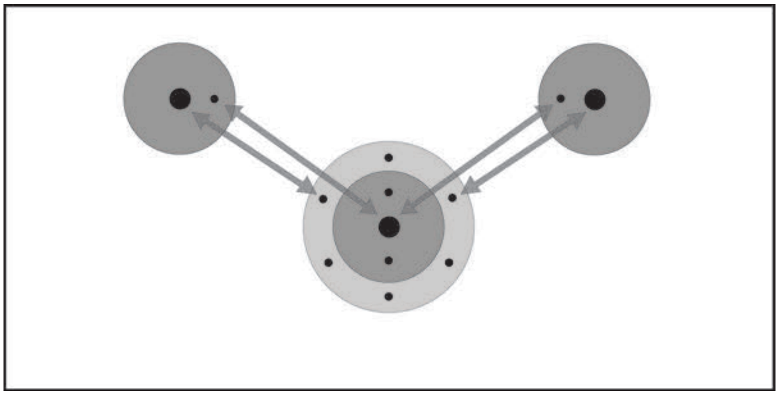
- An electron from a new atom would have to join an atom in the H_2 molecule on the next energy level, further from the nucleus where it would not feel a strong enough attraction.
- An electron from a hydrogen atom already in the H_2 molecule and close to the nucleus would need to move further away to share with the new atom.

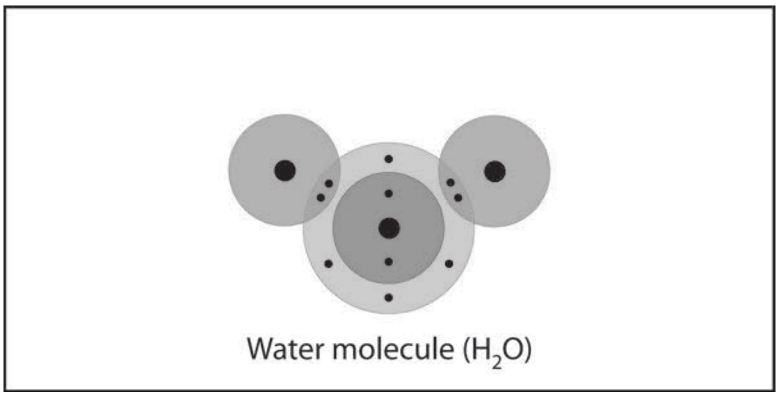
Both of these possibilities would make the molecule less stable and would not happen.

Covalent bonding also happens in a water molecule. When hydrogen atoms and an oxygen atom get close enough together, the electrons from the atoms feel the attraction from the other atom's protons.



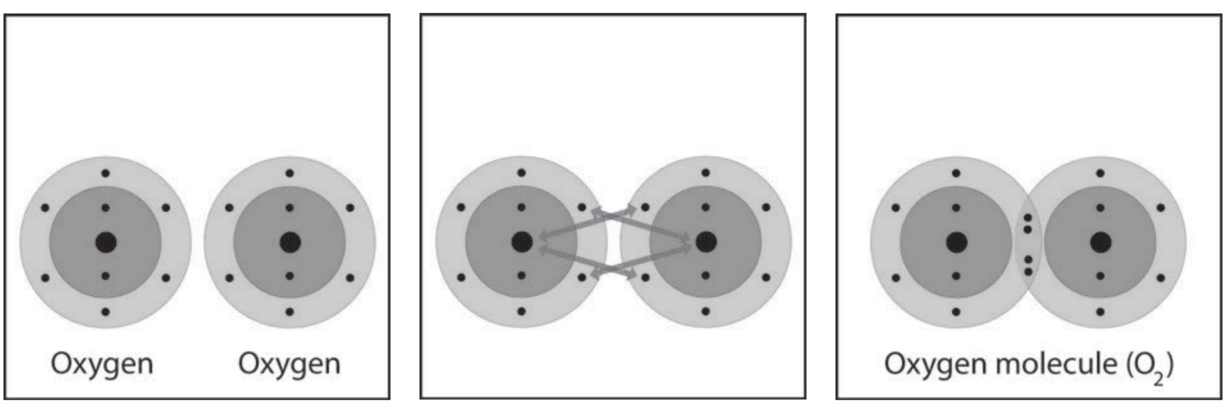
Because there is both a strong enough attraction between the atoms and room for electrons in their outer energy levels, they share electrons. This forms a covalent bond.



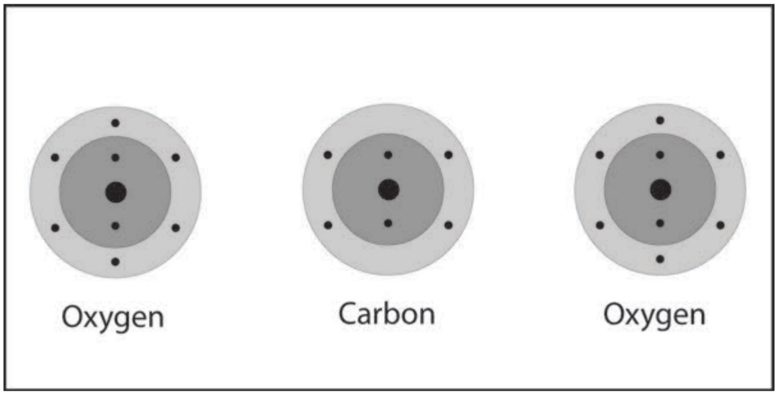


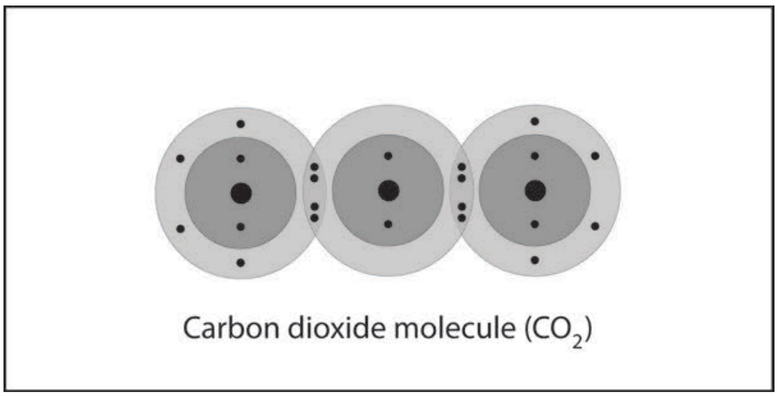
Two oxygen atoms form a double-bond

Oxygen molecules that are present in our air are made up of two oxygen atoms bonded together. Each oxygen atom has 6 valence electrons. When oxygen atoms get close together, the attractions from the nucleus of both atoms attract the outer electrons of the other atom. In this case, 2 electrons from each atom are shared. This is called a double bond.



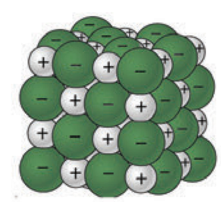
A carbon atom and two oxygen atoms bond to make carbon dioxide (CO_2)





Ionic bonding

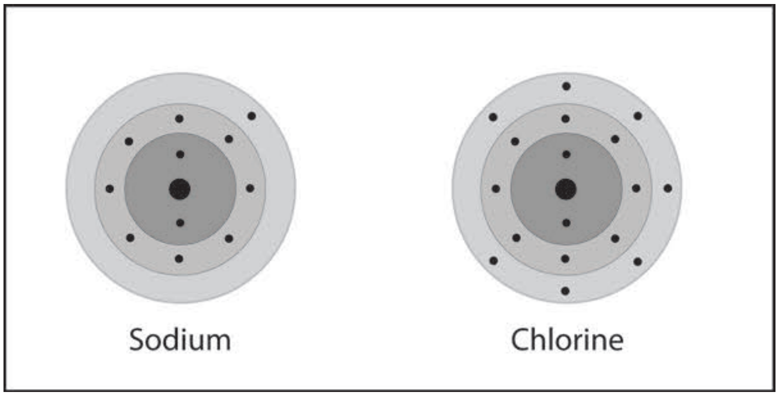
There is another type of bond called an *ionic bond*. One of the most common substances formed by ionic bonding is salt or sodium chloride (NaCl). Look at the model of sodium chloride. The spheres with the "+" and "–" signs on them are called *ions*.

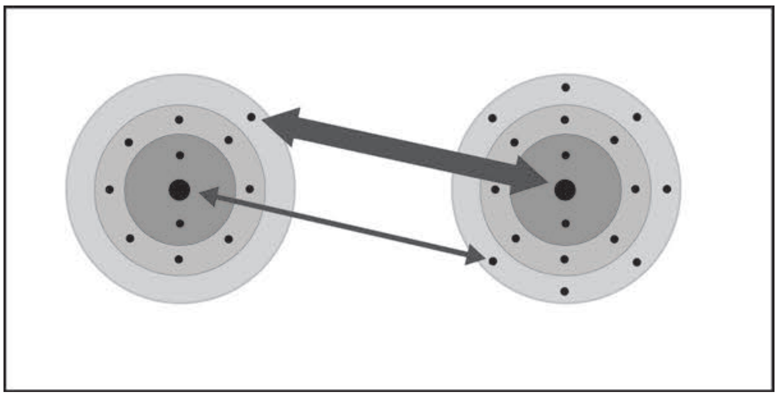


The larger green ones are chloride ions and the smaller gray ones are sodium ions. These ions are formed from chlorine and sodium atoms.

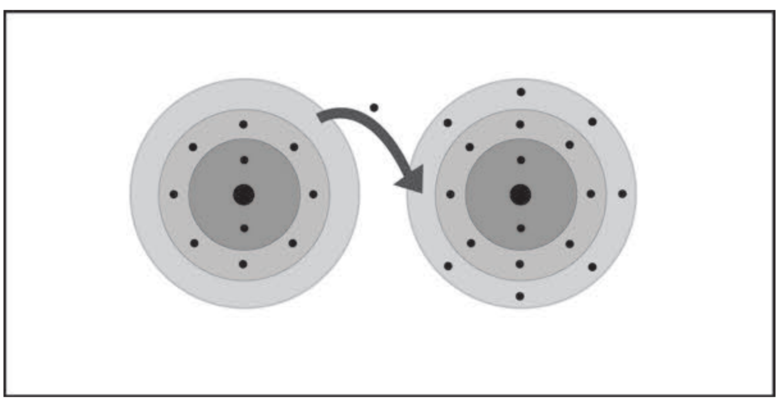
When a sodium and chlorine atom get close enough together, the electrons from the atoms feel the attraction of the protons in the nucleus of the *other* atom.

Chlorine has a stronger attraction for electrons than sodium (shown by the thicker arrow).



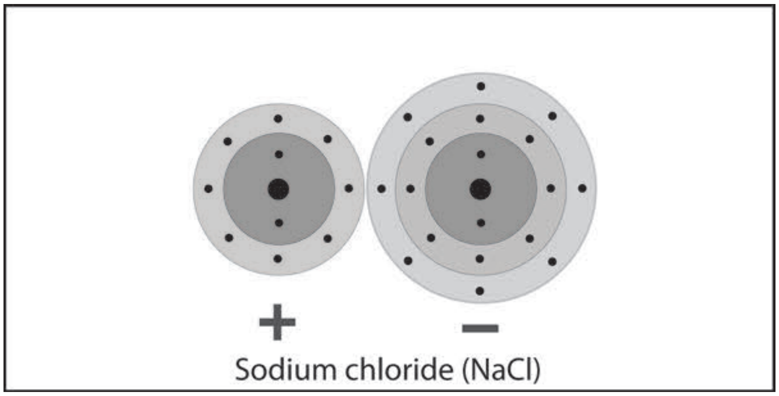


During the interactions between the atoms, the electron in sodium's outer energy level is transferred to the outer energy level of the chlorine atom.



Chlorine *gains* an electron so that the chloride ion has 18 electrons and 17 protons. Since the chloride ion has one more electron than proton, chloride is a *negative* ion with a charge of -1. Sodium loses an electron leaving it with only 10 electrons but 11 protons. This makes sodium a *positive* ion with a charge of +1.

Oppositely charged ions attract each other forming an ionic bond. The bonded ions are more stable than the individual atoms were.

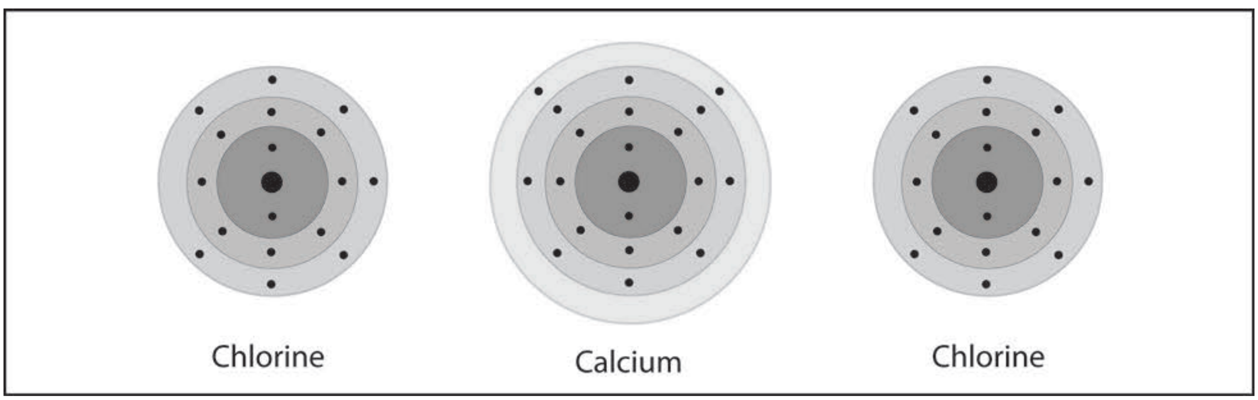


When ions form, atoms gain or lose electrons until their outer energy level is full.

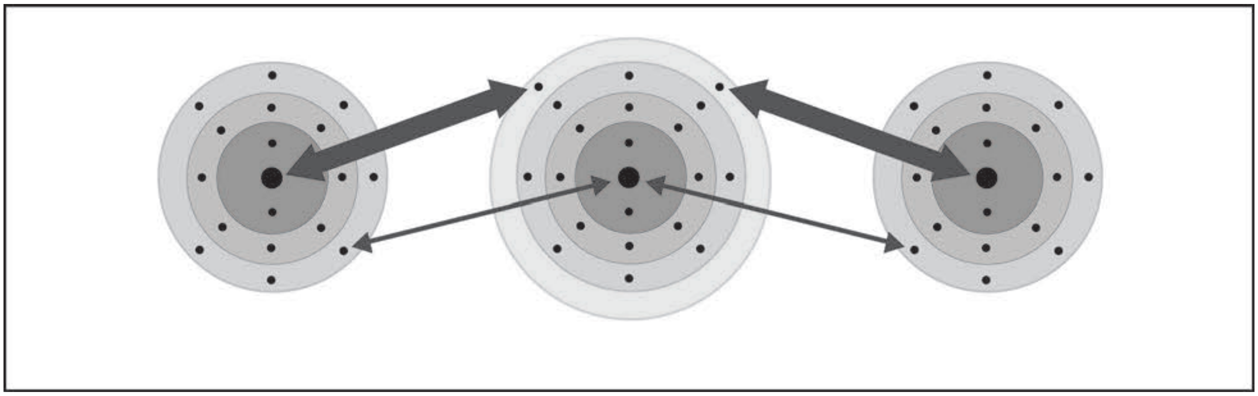
For example, when sodium loses its one outer electron from the third energy level, the second level becomes the new outer energy level and is full. Since these electrons are closer to the nucleus, they are more tightly held and will not leave.

When chlorine gains an electron its third energy level becomes full. An additional electron cannot join because it would need to come in at the fourth energy level. This far from the nucleus, the electron would not feel enough attraction from the protons to be stable.

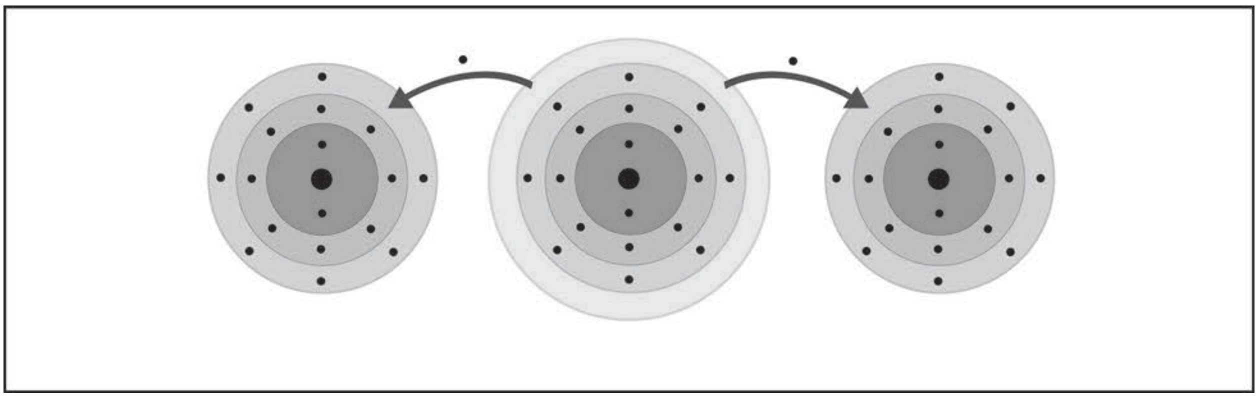
Ionic bonding in calcium chloride ($CaCl_2$)



The protons of the calcium atom attract the electrons from the chlorine atom. The protons of the two chlorine atoms attract the electrons from the calcium atom more strongly as shown by the thicker arrows.

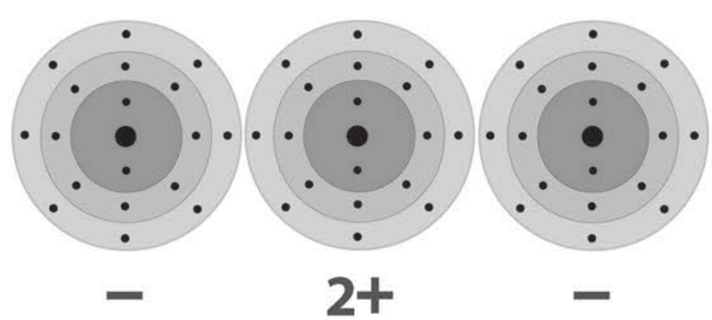


During the interactions between the atoms, the two electrons in calcium's outer energy level are transferred to the outer energy level of each of the chlorine atoms.



Each chlorine atom *gains* an electron so that the chloride ion has 18 electrons and 17 protons. This makes each chloride a *negative* ion with a charge of –1. Calcium *loses* two electrons leaving it with only 18 electrons and 20 protons. This is makes calcium a *positive* ion with a charge of +2.

Oppositely charged ions attract each other forming an ionic bond. The bonded ions are more stable than the individual atoms were.



Calcium chloride ($CaCl_2$)

Chapter 5, Lesson 1: Water is a Polar Molecule

Key Concepts

- The water molecule, as a whole, has 10 protons and 10 electrons, so it is neutral.
- In a water molecule, the oxygen atom and hydrogen atoms share electrons in covalent bonds, but the sharing is not equal.
- In the covalent bond between oxygen and hydrogen, the oxygen atom attracts electrons a bit more strongly than the hydrogen atoms.
- The unequal sharing of electrons gives the water molecule a slight negative charge near its oxygen atom and a slight positive charge near its hydrogen atoms.
- When a neutral molecule has a positive area at one end and a negative area at the other, it is a *polar* molecule.
- Water molecules attract one another based on the attraction between the positive end of one water molecule and the negative end of another.

Summary

Students will be introduced to the idea that water has a slight positive charge at one end of the molecule and a slight negative charge at the other (a polar molecule). Students view animations, make illustrations, and use their own water molecule models to develop an understanding of how the polar nature of water molecules can help explain some important characteristics of water.

Objective

Students will be able to explain, on the molecular level, what makes water a polar molecule. Students will also be able to show in a drawing that the polar nature of water can explain some of water's interesting characteristics and help explain its evaporation rate compared to a less polar liquid.

Evaluation

The activity sheet will serve as the "Evaluate" component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles. Isopropyl alcohol is flammable. Keep it away from flames or spark sources. Read and follow all warnings on the label. Use in well-ventilated room. Dispose of small amounts down the drain or according to local regulations. Have students wash hands after the activity.

Materials for Each Group

- Styrofoam water molecule models from Chapter 2, Lesson 2 (two per student)
- Permanent markers (blue and red)
- Isopropyl alcohol (70% or higher)
- Water
- Brown paper towel
- Droppers

Note about the Materials

Students made molecular models of the water molecule using Styrofoam balls and toothpicks in Chapter 2, Lesson 2. Give each student two of these water molecule models for this activity.

ENGAGE

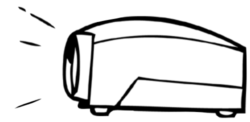
1. Show students examples of water molecules' attraction for one another.

Remind students that in Chapters 1 and 2, they investigated the behavior of water at different temperatures and explored the state changes of water. Many of the explanations were based on the idea that water molecules are attracted to one another. Remind students that in Chapter 4 they looked at the covalent bonding between oxygen and hydrogen, which creates the water molecule. Now students will look more closely at the details of the covalent bonds in a water molecule to understand why water molecules are attracted to one another.

Project the video *Water Balloon*.
www.middleschoolchemistry.com/multimedia/chapter1/lesson1#water_balloon
This video was shown in Chapter 1, Lesson 1 to show that water molecules are attracted to one another.

Project the video *Water Fountain*.
www.middleschoolchemistry.com/multimedia/chapter5/lesson1#water_fountain
Point out that the water is able to stay together in these arcs because water molecules are very attracted to each other.

EXPLAIN



2. Show molecular model animations that illustrate why water molecules are attracted to each other.

Project the animation *Polar Water Molecule*.
www.middleschoolchemistry.com/multimedia/chapter5/lesson1#polar_water_molecule

First Frame of the Animation

- **Electrons are shared between atoms in a covalent bond.**
  Remind students how the shared electrons in a water molecule are attracted to the protons in both the oxygen and the hydrogen atoms. These attractions hold the atoms together.
- **Water molecules are neutral.**
  Be sure students realize that no protons or electrons are gained or lost. The water molecule has a total of 10 protons and 10 electrons (8 from the oxygen atom and 1 from each of the two hydrogen atoms). Since it has the same number of protons and electrons, the water molecule is neutral.

Click "Play"

- **The electron cloud model shows where electrons are in a molecule.**
  Tell students that another way to see the difference in where the electrons are is by using the electron cloud model. Remind students that it's impossible to know the exact location of an electron, so sometimes the regions occupied by electrons are shown as "clouds" around the nucleus in an atom or molecule.
- **Unequal sharing of electrons makes water a polar molecule.**
  Tell students that the oxygen atom attracts electrons a little more strongly than hydrogen does. So even though the electrons from each atom are attracted by both the oxygen and the hydrogen, the electrons are a bit more attracted to the oxygen. This means that electrons spend a bit more time at the oxygen end of the molecule. This makes the oxygen end of the molecule slightly negative. Since the electrons are not near the hydrogen end as much, that end is slightly positive. When a covalently bonded molecule has more electrons in one area than another, it is called a *polar* molecule.
- **The electron cloud model can show an unequal sharing of electrons.**
  Point out that the electron cloud around the oxygen is darker than the electron cloud around the hydrogen. This shows that electrons are more attracted to the oxygen end of the molecule than the hydrogen end, making the water molecule polar.

Click "Next"

- **Color can be added to an electron cloud model to show where electrons are more or less likely to be.**
  Tell students that this is another model of a water molecule. In this model, color is used to show the polar areas of the water molecule. The negative area near the oxygen atom is red, and the positive area near the hydrogen atoms is blue.

Project the animation *Attraction between water molecules*.
www.middleschoolchemistry.com/multimedia/chapter5/lesson1#attraction

Ask students:

- **What do you notice about the way water molecules orient themselves?**
  The red (oxygen) area of one water molecule is near the blue (hydrogen) end of another water molecule.
- **Why do water molecules attract one another like this?**
  Since the oxygen end of a water molecule is slightly negative and the hydrogen end is slightly positive, it makes sense that water molecules attract one another.

Give each student an activity sheet.
Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.

3. Show students that the bonds between atoms in a molecule are different from the polar attractions between molecules.

Project the image *Attractions on different levels*
www.middleschoolchemistry.com/chapter5/lesson1#attractions

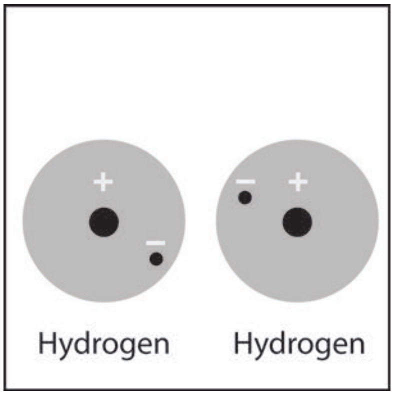
Students may be confused about the bonds within a water molecule and the attractions between water molecules.

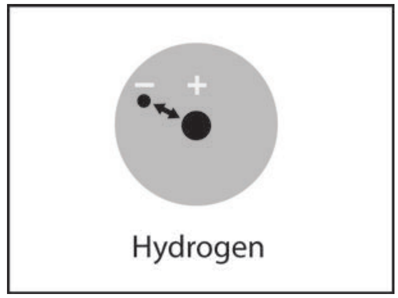
The bonds *within* molecules and the polar attractions *between* molecules
Explain to students that the interaction between the oxygen of one water molecule and the hydrogen of another is different than the sharing of electrons between the oxygen and the hydrogens within the water molecule itself.

It's all about attractions between positive and negative.
Point out to students that attractions between positive and negative works on three differ-

ent levels.

1. A single *atom* stays together because of the attraction between the positively charged protons and the negatively charged electrons.





2. In a molecule, *two or more atoms* stay together because of the mutual attraction between the positively charged protons from one atom and the negatively charged electrons from the other atom. This causes the covalent or ionic bonding that holds atoms or ions together.