Atomic Radii

Look at Figure 2. It shows the relative atomic radii for the first 89 elements (except for the actinide and lanthanide series elements). The radii are not all the same, are they? It probably shouldn’t surprise you, knowing all you do about atoms, that they come in a range of sizes. It’s not a huge range, however. The smallest is hydrogen, which has a radius of about 1/3 Ångström (Å), and the largest (among known atoms) is Francium, which has a radius that is estimated at 2.8 Å. So the radius of the largest neutral atom is less than seven times larger than that of the smallest. All but 13 of the atoms illustrated have radii between 1 and 2 Å. Those having radii smaller than 1 Å have a green fence drawn around them, while those having radii greater than 2 Å are enclosed in a purple fence.

Before we go on, there is something you should be reminded about. Although the radius is a good tool for describing the size of something that is approximately spherical, it can be a little misleading. That’s because a twofold increase in radius has an eightfold effect on volume. So an atom that has a radius of 2 Å takes up eight times the volume of an atom having a radius of 1 Å.

Figure 3 shows the A group elements separately, because the trends among atoms are more obvious if we take out the transition elements. (Adding electrons to the d sublevel doesn’t have as much impact on the dimensions of an atom as adding electrons to s and p sublevels. Adding electrons to the f sublevel—in the lanthanide and actinide series—will have even less impact.) Looking at the Figure, then, notice that potassium (group IA, period 4) is clearly larger than lithium (group IA, period 2). This is precisely what you would expect, since period 2 atoms have electrons in only two principal energy levels, while period 4 atoms have electrons in four. The \( n = 4 \) energy level is higher in energy and farther away from the nucleus than \( n = 2 \). Since the size of an atom is determined by its outermost electrons, we would expect the potassium atom to have a larger radius than the lithium atom. In fact, as we move down any given group, we are filling higher and higher energy levels with electrons, so the atomic radius increases.

Figure 2. These are the relative atomic radii (in Å) of the neutral atoms. Except for a few large atoms and a few small ones, most of them fall within a pretty limited size range. As a point of reference, for comparison, chlorine (group VIIA, period 3) has a radius of about 1 Å.
Atomic radius also changes as you move along a period. As you move from left to right within a given period the atomic radii decrease. That's because, as you move from left to right, the number of positively-charged protons in the nucleus increases. Therefore, the electrons, attracted by this increasing charge concentration, will be held more and more tightly.

Overall, notice how the atomic radii increase as you near the lower left corner of the table (francium) and how they decrease as you move upward and to the right (toward helium). That is, with the most notable exception of hydrogen, which has the smallest of all atomic radii.

Ionic Radii

Would you expect an ion, having either one more or one less electron than its neutral atom, to occupy exactly the same amount of space as the neutral atom itself? Of course not! Direct your attention to Figure 4, which compares the radii of the A-group elements. The noble gases were left off because they form ions only rarely and with difficulty. First, notice that, as you go down a group, the trend is the same for both positive ions and negative ions as it was for the neutral elements—the radius increases. It does so for the same reason—there are more and more electrons filling higher and higher energy levels.

However, notice some differences. First of all, notice that when an element forms a cation (shown in red), its radius decreases. In fact, the cation of an element will always be smaller than the element itself. Why? You could answer this question yourself—to form a cation, one or more electrons must be removed, while the number of protons remains constant. Therefore, the same positive charge that was attracting a certain number of electrons in the neutral element is now pulling on fewer negative charges in the ion. The fewer electrons that remain can be held more tightly. Thus the radii decrease.

The same logic applies to anions (shown in yellow), but the effect is reversed. They have gained one or more extra electrons. What effect should this have on the radius? Well, again the number of protons remains the same, so the same number of positively-charged particles is now pulling on more negatively-charged ones. This means that the electrons cannot be held quite as tightly. That extra electron will also assume a high energy level. Both of these factors tend to increase the radius of the ion. The anion of an element will always have a larger radius than the corresponding neutral element.

If we consider only those elements that form cations (shown in red), then the trend as we move from left to right within a period is the same as for the neutral elements—the radius decreases. Remember, though, as you move from left to right within a period that you are forming cations by removing more and more electrons. Sodium loses one electron to form Na⁺;
upper right. That is, of course, with the notable exception of hydrogen. It has only one electron. When that electron is removed, its radius becomes vanishingly small—the radius of a single proton.

Enough about cations. Is there a trend among those elements that form anions? Yes. First of all, notice that the elements on the left of the table will not gain electrons. You must be in group IVA or above to have any desire to add an electron to your quiver. That is, with the notable exception of hydrogen, which has only one electron in its 1s orbital. It will form an anion.

As was the case with cations, we have to remember that some of the ions shown on the table represent elements that are gaining more than one electron. So fluorine gains only one electron to form $\text{F}^-$; oxygen gains two electrons to form $\text{O}^{2-}$; and nitrogen gains three electrons to form $\text{N}^{3-}$. So the anions should get looser as you move from right to left—fewer protons—looser, more energetic electrons. And this is generally what we see (except for hydrogen—it has no anions in its period with which to compare it.)

Just for the fun of it, we’ll now put the transition elements back into the Figure to see what effect they have on the trends. This is done in Figure 5. Now you can see what I meant when I said that adding $d$-sublevel electrons doesn’t have quite the same amount of impact. There are some interesting atoms and ions that stand out here that we could talk about individually (although we won’t), but generally the transition metals have a similar look about them: (1) their radii are all in the range from 1.24 Å (Ni) to 2.0 Å (Ac); (2) none of them form anions; and (3) the ions they form are a little more than half the size (radius) of the corresponding neutral atoms.

<table>
<thead>
<tr>
<th>IA</th>
<th>IIA</th>
<th>IIIA</th>
<th>IVA</th>
<th>VA</th>
<th>VIA</th>
<th>VIIA</th>
</tr>
</thead>
<tbody>
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</table>

Figure 4. Positive ions (red) are smaller than the neutral atoms that form them (blue), because the excess positive charge holds the remaining electrons more tightly. Negative ions (yellow) are larger than their neutral atoms, because they have an extra electron relative to their numbers of protons. The trends in the radii as you move down a group or across a period are the same for the ions as for the neutral atoms—smaller from left to right, larger from top to bottom.

Magnesium loses two electrons to form $\text{Mg}^{2+}$; and aluminum loses three electrons to form $\text{Al}^{3+}$. So the cations should get tighter and tighter as you proceed to the right. This is precisely what we see in the Figure. From IA to IIA to IIIA, there are more and more protons holding onto fewer and fewer electrons—the radii decrease rapidly. Because of this trend, it gets harder and harder to form positive ions as you move to the right. You’ll notice that nitrogen and fluorine have no positive ions reported in the Figure. They won’t give up those electrons.

So then, the cations will be seen to decrease in size, as you move from the lower left of the table to the
Ionization Energy

Understanding the trends in atomic radius makes it easier to understand other atomic properties of the elements. For example, consider the ionization energy of an element. Ionization energy is defined as the amount of energy required to remove an electron from an atom to form a cation. Why do ionization energies matter? They are the measure of how easily an atom will give up electrons. This will tell us much about its reactivity as we move ahead.

As an example, consider the formation of the lithium cation:

\[ \text{Li} + \text{energy} \rightarrow \text{Li}^+ + e^- \]

The energy required to remove an electron from an atom is related to the energy the electron already has. If an electron has a great deal of energy to start with, relatively little is required to remove it from its nucleus.

As you would expect, an electron having relatively high energy and having a location farther from the positive charge of the nucleus would require less energy to remove it. An atom having such an electron would have a lower ionization energy than one having a smaller radius. Look at Figure 6, which shows the first ionization energies of the group IA elements. (The first ionization energy is the energy required to remove the outermost electron.) In group IA, you can see that the ionization energies decrease as you move down the table. This is true for all the group A elements.

Hydrogen also follows this trend except that, because it has only one electron, and it is in the 1s sub-
level, it is the hardest one to remove. Its ionization energy is far greater than any other element in its period—it’s way out of line. Our figure shows only first ionization energies. Every element except hydrogen has a series of ionization energies because they all have more than one electron that may be removed and each of as you go across. When the d sublevels are full (Zn, Cd, Hg), there is a noticeable decrease in ionization energy as you jump into group IIA (Ga, In, Tl). Of course, this makes sense. You might expect it to be a little harder to remove electrons from a full sublevel than from one that contains only a single electron.

What about any trend across the periods? Well, what was the trend among the atomic radii? As you moved from left to right, the electrons were held more tightly because of a more concentrated positive charge in the nucleus. If electrons are held more tightly, would you expect to need more energy or less to remove one? I hope you said, “More.” Therefore, as you move from left to right within a period, ionization energy tends to increase. You need to be aware, though, that this trend is not as consistent as it is within groups. As you can see in the Figure, beryllium and magnesium are the major trend breakers among the group A elements. As usual, hydrogen is in a class by itself.

Now, just for the fun of it, let’s put the transition elements back in and see their effects on these trends. (See Figure 7.) The ionization energies of the transition elements don’t change much as you go down the table, and they increase very gradually.
Properties of Matter

13: With the Exception of Hydrogen

Figure 6. These are the ionization energies (in "electron volts") of all the elements. There is a much steadier trend among the metals in this table than we have seen before. As a rule, electrons get more and more difficult to remove as you move across the table from left to right. However, as you jump from the transition metals to the post-transition elements ionization energies take a dip—that d sublevel is full at the end of the transition metals, and happy to stay that way, thank you very much! You see the same effect (to a lesser extent) in moving from the alkaline earth metals to the transition metals. Once the s sublevel is filled, the next added electron will be slightly easier to remove.

Exercises

1) Which element has the highest ionization energy? Why?

2) Which has the lowest ionization energy? Why?

3) Which neutral element has the greatest atomic radius? Why?

4) Which has the smallest atomic radius? Why?

5) Which cation has the greatest radius? Why?

6) Which cation has the smallest radius? Why?

7) Which anion has the largest radius? Why?

8) Which anion has the smallest radius? Why?

9) Which has the smaller radius, Cl or Cl\(^{-}\)? Why?

10) Which has the greater ionization energy, Mg or Mg\(^{2+}\)? Why?
14: Quiz 2

1) Be able to identify the crystalline and amorphous substances mentioned.
2) Understand the difference between physical and chemical properties and processes.
3) Understand the difference between homogeneous and heterogeneous substances.
4) Understand how properties of substances are used in separating mixtures.
5) Be able to perform specific heat calculations.
6) Know the order in which orbitals fill as the number of electrons increases from element to element. You may do this by being able to reproduce the chart that is found on the “electron configuration” sheet in your lab set.
7) Understand the meanings of the periods and groups in the periodic table in terms of the filling of orbitals (as shown in the periodic table inserted in Lesson 13).
8) Know the informal names of the element groups and series. They are illustrated in a periodic table in Lesson 8 (alkali metals, alkaline earth metals, etc.).
9) As in every other lesson (whether or not we say) know and understand anything in bold or italic typeface.
15: ISOTOPES AND ATOMIC MASS

Isotopes

In previous lessons we have pointed out that the atomic number of an element always tells us the number of protons in the nucleus of an atom of that element. However, all atoms of a given element may not have the same mass. Why? Remember we said that the mass of an atom is essentially the mass of its nucleus. So if the mass of one atom of an element is different from the mass of another atom of the same element, and yet they both contain the same number of protons, then it must necessarily follow that the number of neutrons is different in the two atoms. Atoms which contain the same number of protons but have different masses are called isotopes. For example, there are three known isotopes of hydrogen:

protium (1 proton, 0 neutrons) - “normal” hydrogen
deuterium (1 proton, 1 neutron)
tritium (1 proton, 2 neutrons)

Notice that each of these isotopes has the same atomic number—the number which corresponds to the number of protons. Since there is only one proton, there is only one positive charge in the nucleus. In order to be charge balanced, there must be only one electron present in each of the isotopes. But because of the added neutrons, the mass of deuterium is very close to twice that of protium, and the mass of tritium is very close to 3 times that of protium.

Because we occasionally need to communicate the different isotopes of the elements, we need a method to distinguish them. Isotope notation was invented for that purpose. The general notation is given by:

where:
X = the element’s symbol
A = mass number (number of protons + neutrons)
Z = atomic number

Here is an example of isotope notation:

\[ ^{25}_{11}\text{Na} \]

To help you remember which letter corresponds to the mass number and which corresponds to the atomic number, just remember that there may be many mass numbers (isotopes) for a single element, so the one you are writing is just “\( A \) mass number.” But there is only one atomic number for any element, so the one you are writing is “\( Z \) only atomic number.” Always put the big number on top. It’ll be the mass number. Here are isotope notations for the three isotopes of carbon:

\[ ^{12}_6\text{C} \quad ^{13}_6\text{C} \quad ^{14}_6\text{C} \]

Sometimes, since the atomic number is often known or so easily obtained from the periodic table, the \( Z \) is omitted. So the notation for carbon-14 would simply look like this:

\[ ^{14}_6\text{C} \]

Also, if you are given the notation for an isotope, you should be able to figure out how many protons, electrons and neutrons are present. For example, if you were asked how many protons, electrons and neutrons are in the isotope of fluorine represented by \( ^{19}_9\text{F} \), you would first look up \( \text{F} \) in the periodic table to see that the atomic number of fluorine is 9. Therefore, there are 9 protons. Since it is electrically neutral, there must also be 9 elec-
trons. Finally, because the mass number is 19, the number of neutrons can be determined by subtracting the number of protons from it. There are $19 - 9 = 10$ neutrons.

All the common elements have known isotopes; several elements have as many as 30 to 50 isotopes each. However, only a few of these will account for more than one percent of the element found in nature. As in the case of zinc, there are even some elements of which less than half of the naturally occurring element is in the form of its most prominent isotope. The rest is made up from small amounts of each of a multitude of isotopes.

**Atomic Mass**

As we mentioned before, individual atoms have very small masses (remember the mass of a hydrogen atom was $1.673 \times 10^{-24}$ g). Therefore, it is simply not practical to compare the actual masses of various atoms when the numbers are so small. For this reason a relative scale was devised so that the masses of different elements can be more conveniently compared. Carbon-12 ($^{12}$C) was chosen as the standard against which everything else would be measured, and it was assigned a mass of 12 atomic mass units (amu or AMU). Therefore, one amu equals $1/12$ the mass of a $^{12}$C atom. Just be aware that when you read atomic mass values for a given element on the periodic table, these values were calculated relative to $^{12}$C.

Even though the scale of atomic mass is built around $^{12}$C, natural carbon is not exclusively made up of $^{12}$C. It also has a little bit of $^{13}$C in it. If we want to do chemistry in the real world we won’t be doing chemistry with pure $^{12}$C, but with a mixture of $^{12}$C and $^{13}$C! (See Figure 1.) We are not so much concerned with the atomic mass of $^{12}$C, but rather the average atomic mass of all carbon. There are two ways we could calculate this average atomic mass.

Look now at Figure 2. The fishing sinkers shown here are of two different masses. What is the average mass of all the sinkers? Well, you could figure it out the way you have always calculated averages. That is, by simply adding the
This second method of calculating the average is called the **weighted average** method. The individual masses are “weighted” by their corresponding proportions to calculate an average. Look at your periodic table and find the atomic mass of carbon. It is given as 12.01, not simply as 12. Why isn’t the atomic mass of elemental carbon recorded as 12? Because carbon doesn’t exist in nature as only $^{12}\text{C}$, but as a mixture of 98.89% $^{12}\text{C}$, and 1.11% $^{13}\text{C}$. So the mass given in the periodic table is a weighted average of these two isotopes.

You would calculate the weighted average mass of carbon like this. $^{12}\text{C}$ has a mass of 12 amu by definition. But how much importance or “weight” should we give to it in calculating the average atomic mass? Well, it accounts for 98.89% of all the mass, so it should get 0.9889 times the total importance in determining the average atomic mass. $^{13}\text{C}$ has a known mass of 13.0033 amu (to six significant digits) but it accounts for only 1.101% of the total mass of carbon. It should get only 0.01101 times the total importance in our calculation. So we should *weight* each atomic mass by its relative abundance to get its “weighted mass.” The sum of the weighted masses is the average mass.

Here it is, plain and simple:

\[
\text{Average mass (amu)} = 0.9889 (12 \text{ amu}) + 0.01110 (13.0033 \text{ amu}) = 12.01 \text{ g}
\]

(Remember, the 12 is a number that is known by definition, so its precision is infinite. If you need help remembering not to let it limit the precision of your calculations, underline it when you write it down, like I did.)

All of the atomic masses of the elements that appear in the periodic table are the weighted averages of the isotopes that make them up. If you know the relative abundances of all the naturally-occurring isotopes of an element and the atomic mass of each isotope, you can calculate the average atomic mass for that element.

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**Properties of Matter**

masses of the individual sinkers and dividing by the total number of sinkers like this:

<table>
<thead>
<tr>
<th>Sinker</th>
<th>Mass</th>
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<tbody>
<tr>
<td>1</td>
<td>4 g</td>
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<tr>
<td>2</td>
<td>4 g</td>
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<td>3</td>
<td>4 g</td>
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<td>6</td>
<td>4 g</td>
</tr>
<tr>
<td>7</td>
<td>4 g</td>
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<tr>
<td>8</td>
<td>10 g</td>
</tr>
<tr>
<td>9</td>
<td>10 g</td>
</tr>
<tr>
<td>10</td>
<td>10 g</td>
</tr>
</tbody>
</table>

**Total** 58 g

Average mass (g) = \[\frac{58 \text{ g}}{10} = 5.8 \text{ g}\]

This is a little tougher to do with atoms than it is with sinkers (unless you are asked to calculate the average mass of a trillion sinkers.)

The other way we could do it is by multiplying each of those masses by the portion of all sinkers it represents. In this case 7/10 of the sinkers were 4-g sinkers and 3/10 were 10-g sinkers. So:

Average mass (g) = \[0.7 \times 4 \text{ g} + 0.3 \times 10 \text{ g} = 5.8 \text{ g}\]

Both methods give the same answer. When working with atoms, the first method is not an option. There are few circumstances under which we would ever count and weigh individual atoms. (We’ll tell you how this can be done later.) Using the second method you needn’t count and weigh individual atoms. You just need to know the masses of the individual atoms that make up the two different kinds, and the proportion of the entire mass made up by each.

To weigh out one trillion sinkers would require 10,000 workers, working 8 hours per day, each weighing one sinker per second, over 13.35 years to complete the task, not including restroom breaks.
Now it's your turn!

Exercises

1) For the following isotopes, give the number of protons, electrons and neutrons.
   a) $^{238}_{92}\text{U}$
   b) $^{81}_{35}\text{Br}$
   c) $^{63}_{29}\text{Cu}$

2) Complete the following table:

<table>
<thead>
<tr>
<th>isotope notation</th>
<th>A</th>
<th>Z</th>
<th>no. of protons</th>
<th>no. of electrons</th>
<th>no. of neutrons</th>
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</thead>
<tbody>
<tr>
<td>$^{79}\text{Br}$</td>
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<td>90</td>
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<td>$^{1}\text{N}$</td>
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3) Magnesium occurs in nature as a mixture of three isotopes, $^{24}$Mg (23.98504 amu), $^{25}$Mg (24.98584 amu), and $^{26}$Mg (25.98259). Their relative abundances are 78.70%, 10.13%, and 11.17%, respectively. Calculate the average atomic mass of Mg.

4) There are only two naturally-occurring isotopes of iridium, $^{191}$Ir and $^{193}$Ir. $^{191}$Ir has a mass of 190.9609 amu and $^{193}$Ir has a mass of 192.9633 amu. Their relative abundances are 37.30% and 62.70%, respectively. Calculate the average atomic mass of Ir.