
Welcome!

This is a chemistry lesson, covering *atomic mass*.

I will guide you step-by-step.

I will be asking you many questions along the way.

Each time I ask a question, **you should attempt to answer the question on your own** before you scroll down to view my answer.

Feel free to ask me questions at any point.

When you're done with your questions, say "Return to the script, at the exact same page we left off."

This is a lesson in the chapter "Atoms, Molecules, and Compounds",
which is the second chapter of the course, "Chemistry, Explained Step by Step".

In some textbooks, the material in this lesson
is covered in the chapter on Stoichiometry,
rather than in the chapter on Atoms, Molecules, and Compounds.
If your class hasn't covered atomic mass yet, you may want to postpone this lesson.

This lesson primarily builds on the material we covered in the previous lesson on Isotopes,
as well as on earlier lessons on Atomic Number and Mass Number
and on Atoms, Protons, Neutrons, and Electrons.

You should complete those lessons before working through this one.

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Let's review a mathematical technique that we will need during this lesson.

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1. Suppose that, during your academic career, you receive 30% A's, 60% B's, and 10% C's. Calculate your grade point average.

Answer:
Your grade point average was 3.2

Solution:
? = grade point average

I think you know that, when calculating your GPA,
A's are worth 4 points,
B's are worth 3 points, and
C's are worth 2 points.

We need to convert the percentages into decimals.
We do this by moving each decimal point two places to the left.
30% = .3
60% = .6
10% = .1

Now, we use the mathematical technique of calculating a *weighted average*.
grade point average = .3(4) + .6(3) + .1(2)
So, grade point average = 3.2

The numbers .3, .6, and .1 are referred to as the *weights*.

I hope you agree that the weights for a weighted average should add up to 1 (or, 100%).

Check:
sum of weights for last problem = .3 + .6 + .1
So, sum of weights = 1
Yes, the weights add up properly.

When working with experimental data,
it is sufficient for the weights to add up to *approximately* 1 (or, approximately 100%).

Does our answer for problem 1 make sense?

Most of the weight (.6) is on B’s, so we expect our answer to be close to 3.

More weight goes to A’s (.3) than to C’s (.1),
so we expect our answer to be slightly bigger than 3,
rather than slightly smaller than 3.

Our answer was 3.2, which is slightly bigger than 3,
so our answer does make sense.

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2. True or False?
In the previous problem, there is no individual class for which you received 3.2 grade points.

If true, explain why it’s true.
If false, explain why its false.

Answer: True

According to the information in the previous problem,
during the course of your academic career,
you received an A, worth 4 grade points, in 20% of your classes;
you received a B, worth 3 grade points, in 70% of your classes; and
you received a C, worth 2 grade points, in 10% of your classes.

As you can see, there was no individual class for which you received 3.2 grade points.

Rather, your “grade point average” of 3.2 represents
a *weighted average* of the grade points you received in all of your classes.

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3. Every day you buy a sandwich for lunch.
80% of the days, you buy a cheap sandwich for $4.
20% of the days, you buy an expensive sandwich for $10.
What is the average price you pay for your sandwiches?

Answer:
The average price you pay for a sandwich is $5.20.

Solution:
? = average price

We are told the percent of days with two different prices:
80% of the days → $4
20% of the days → $10

We need to convert the percentages into decimals:
80% = 0.8
20% = 0.2

Now, we use the mathematical technique of calculating a *weighted average*.
average price = 0.8($4) + 0.2($10)
So, average price = $5.20

In this example, the numbers 0.8 and 0.2 are referred to as the *weights*.

Check:
Sum of weights = 0.8 + 0.2
So, sum of weights = 1
So, yes, the weights add up properly.

Does our answer make sense?

Most of the weight (0.8) is on the cheaper sandwich, $4,
so we expect the average to be closer to $4 than to $10.

The average is $5.20 — which is closer to $4 than $10 —
so yes, the answer makes sense.

In general, suppose you are dealing with two quantities, Q_1 and Q_2.
Suppose each quantity is associated with a weight, W_1 and W_2, where $W_1 + W_2 = 1$.
Then, the *weighted average* of the quantities = $W_1Q_1 + W_2Q_2$.

Or, suppose you are dealing with three quantities, Q_1, Q_2. and Q_3.
Suppose each quantity is associated with a weight, W_1, W_2, or W_3, where $W_1 + W_2 + W_3 = 1$.
Then, the weighted average of the quantities = $W_1Q_1 + W_2Q_2 + W_3Q_3$.

Similar definitions can be given for the weighted average of four quantities, or five quantities, etc.

The *mass* of an object is defined as
the *quantity of matter* contained in that object.

An elephant has a greater mass than a mouse,
because the elephant contains a greater quantity of matter than the mouse does.

A bowling ball has a greater mass than a nerf ball,
because the bowling ball contains a greater quantity of matter than the nerf ball does.

In chemistry, the most common units used for mass are grams (g).

A paperclip might have a mass of roughly 1 g.

Some other common units for mass are kilograms (kg) and milligrams (mg).

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4. Arrange these units from biggest to smallest:
gram, kilogram, milligram

Answer:
From biggest to smallest, the units are
kilogram, gram, milligram

Analysis:
In our earlier lesson on Unit Conversion and Metric Prefixes,
we learned that kilo = 10^3,
and milli = 10^{-3}.

Because $10^3 > 1 > 10^{-3}$
we know that a kilogram (10^3 grams) is bigger than a gram (1 gram),
and a gram (1 gram) is bigger than a milligram (10^{-3} grams).

For more information and practice problems for working with metric prefixes such as kilo and milli,
review the lesson on Unit Conversion and Metric Prefixes.

The mass of a proton is 1.673×10^{-24} g.

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5. Is 1.673×10^{-24} an ordinary-sized number, a huge number, or a tiny number?

1.673×10^{-24} is a tiny number.

In ordinary notation, 1.673×10^{-24} g would be written as .00000000000000000000000001673 g.
(If you don't know how to convert from scientific notation to ordinary notation, you should review the lesson on Scientific Notation.)

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6. Why is the mass of a proton such a tiny number when measured in grams?

Grams are a convenient unit for measuring the mass of ordinary-sized objects, such as paperclips.

A proton is a tiny, tiny object—far, far smaller than an ordinary-sized object like a paperclip.

So the mass of a proton is far, far smaller than the mass of an ordinary-sized object like a paperclip.

So it makes good sense that, when measured in grams,
the mass of a proton will be a tiny number.

When working with tiny particles like protons, neutrons, and atoms,
it's often not convenient to have to work with tiny numbers like 1.673×10^{-24}.

To avoid this, chemists have invented a unit of mass that's more appropriate for working with atom-sized particles.

The unit is called the "amu".

1 amu is a tiny, tiny amount of mass,
so amu's are a convenient unit for measuring the mass of roughly atom-sized particles.

The mass of a proton is 1.007 amu.

So it should be apparent to you that
$1.007 \text{ amu} = 1.673 \times 10^{-24} \text{ g}$

The mass of a neutron is 1.009 amu,
so you can see that a neutron has approximately the same mass as a proton.

The mass of an electron is 0.0005 amu,
so you can see that an electron is far less massive than a proton or neutron
(roughly 2000 times less massive).

“amu” stands for “atomic mass unit”

Let’s review.

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7. What is the definition of “mass”?

The mass of an object is defined as
the quantity of matter contained in that object.

(This definition is somewhat vague.
If you take a physics class you will learn a more precise definition of mass.)

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8. What are the units for mass that we have discussed in this lesson?
Arrange the units from biggest to smallest.
Give names and abbreviations.

We have discussed the following units for mass, arranged from biggest to smallest:
kilograms = kg
grams = g
milligrams = mg
atomic mass units = amu

Moving on, recall from the previous lesson that
naturally occurring elements are usually a *mixture* of different isotopes.

For example,
naturally occurring magnesium is a *mixture* of
78.99% magnesium-24,
10.00% magnesium-25, and
11.01% magnesium-26.

One atom of magnesium-24 has a mass of 23.9850 amu. One atom of magnesium-25 has a mass of 24.9858 amu. One atom of magnesium-26 has a mass of 25.9826 amu.

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9. What is the average mass of a magnesium atom in a sample of naturally occurring magnesium?

Answer:
The average mass of a magnesium atom is 24.3055 amu.

Solution:
? = average atomic mass of magnesium

We have been given the following information: Magnesium-24: mass = 23.9850 amu, found in 78.99% of magnesium atoms Magnesium-25: mass = 24.9858 amu, found in 10.00% of magnesium atoms Magnesium-26: mass = 25.9826 amu, found in 11.01% of magnesium atoms
Convert the percentages to decimals,
by moving the decimal points two places to the left.
78.99% = .7899 10.00% = .1000 11.01% = .1101

Now, we calculate the weighted average: average atomic mass = .7899(23.9850) + .1000(24.9858) + .1101(25.9826)
So, average atomic mass = 24.3055 amu
So, the average atomic mass of magnesium is 24.3055 amu.

The numbers .7899, .1000, and .1101 are the weights.
Do these weights make sense?

Check:
Sum of weights = .7899 + .1000 + .1101
So, sum of weights = 1

So, yes, the weights add up properly.

When calculating the average mass of a type of atom,
the weights are referred to as the “abundances”.

isotopic abundance:
the relative proportion of a given isotope of an element,
expressed as a percentage of all isotopes of that element found in nature

The abundance of magnesium-24 is 78.99%, or .7899
The abundance of magnesium-25 is 10.00%, or .1000
The abundance of magnesium-26 is 11.01%, or .1101

Does our answer for problem 9 make sense?

Most of the weight (about 79%) is on magnesium-24.
so we expect the average mass to be not too far from the mass of magnesium-24 (23.9850 amu).

The remaining weight (about 21%) is on values greater than 23.9850 amu (24.9858 amu and 25.9826 amu),
so we expect the average mass to be somewhat greater than 23.9850 amu

The weighted average is 24.3055 amu.
This value matches our expectation;
so, yes, our answer makes sense.

The value 24.3055 amu which we calculated in problem 9
is referred to as the *atomic mass* of magnesium.

The *atomic mass* of an element is defined as
the weighted average of the masses
of the naturally occurring isotopes of that element.

(You can see that a more accurate name for “atomic mass” would be “average atomic mass”,
but for simplicity the name “atomic mass” is standard.)

So, as we saw in problem 9,
the atomic mass of magnesium is calculated as
the weighted average of the masses
of the naturally occurring isotopes of magnesium.

The standard units for atomic mass are amu.

The atomic mass for each element is reported in the periodic table.

The decimal numbers in the periodic table (1.008, 4.003, 6.941, 9.012, etc.)
represent the numerical value of the atomic mass for each element, when measured in amu.

Please look up the atomic mass for magnesium in the periodic table.

You should find that the periodic table reports
an atomic mass for magnesium of 24.31 amu.

In problem 9 we calculated
an atomic mass for magnesium of 24.3055 amu,
which, when rounded to two decimal places, matches the value reported in the periodic table.

10. Silicon has three naturally occurring isotopes.

Isotope	Mass (amu)	Abundance (%)
Silicon-28	27.9769	92.23
Silicon-29	28.9765	4.67
Silicon-30	29.9738	3.10

Use this data to calculate the atomic mass of silicon.

Answer:
The atomic mass of silicon is 28.0855 amu.

Solution:
? = atomic mass of silicon

We have the masses and natural abundances of each isotope:
Silicon-28: mass = 27.9769 amu, abundance = 92.23%
Silicon-29: mass = 28.9765 amu, abundance = 4.67%
Silicon-30: mass = 29.9738 amu, abundance = 3.10%

Convert the percentages to decimals.

Now, we calculate the atomic mass,
which is a weighted average of the isotopic masses:
atomic mass = .9223(27.9769) + .0467(28.9765) + .0310(29.9738)
So, atomic mass = 28.0855 amu

So, the atomic mass of silicon is 28.0855 amu.

The abundances .9223, .0467, and .0310 are our weights.
Do these weights make sense?

Check:
Sum of weights = .9223 + .0467 + .0310
So, sum of weights = 1

So, yes, the weights add up properly.

Does our answer for problem make sense?

Almost all the weight (about 92%) is on silicon-28,
so we expect the average mass to be very close to the mass of silicon-28 (27.9769 amu).

The remaining weight (about 8%) is on values greater than 27.9769 amu (28.9765 amu and 29.9738 amu),
so we expect the average mass to be very close to, but a little bigger than, 27.9769 amu

The weighted average is 28.0855 amu,
which is indeed very close to, but a little bigger than, 27.9769 amu.
This value matches our expectation;
so, yes, our answer makes sense.

What's another way you could have checked whether you got the right answer to problem 10?

I hope you thought to check your answer for problem 10
against the value in the periodic table.

The decimal numbers in the periodic table
represent the atomic mass for each element.

You should find that the periodic table reports
an atomic mass for silicon of 28.09 amu.

In problem 10 we calculated
an atomic mass for silicon of 28.0855 amu,
which, when rounded to two decimal places, matches the value reported in the periodic table.

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11. Gallium has two naturally occurring isotopes, ^{69}Ga (mass = 68.9256 amu, abundance = 60.11%) and ^{71}Ga (mass = 70.9247 amu, abundance = 39.89%). Use this data to calculate the atomic mass of gallium.

Answer:
The atomic mass of gallium is 69.7231 amu.

Solution:
? = atomic mass of gallium

We have the masses and natural abundances of each isotope:
Gallium-69: mass = 68.9256 amu, abundance = 60.11%
Gallium-71: mass = 70.9247 amu, abundance = 39.89%

Convert the percentages to decimals.

Now, we calculate the atomic mass,
which is a weighted average of the isotopic masses:
atomic mass = .6011(68.9256) + .3989(70.9247)
So: atomic mass = 69.7231 amu

So, the atomic mass of gallium is 69.7231 amu.

The abundances .6011 and .3989 are our weights.
Do these weights make sense?

Check:
Sum of weights = .6011 + .3989
So, sum of weights = 1
So, yes, the weights add up properly.

Does our answer for problem 11 make sense?

The atomic mass is a weighted average of 68.9256 amu and 70.9247 amu,
so we expect the result to be *between* 68.9256 amu and 70.9247 amu.

Slightly more weight is put on 68.9256 amu (about 60%) than on 70.9247 (about 40%),
so we expect the result to be slightly closer to 68.9256 amu than to 70.9247 amu.

To simplify, we expect our result
to be between 69 and 71,
but slightly closer to 69 than to 71;
i.e., we expect our result to be a little less than 70.

Our result for the weighted average was 69.7231 amu.
This value is indeed slightly closer to 69 than to 71.
So, yes, our answer makes sense.

Hopefully you also checked your answer for problem 11
against the value in the periodic table.

You should find that the periodic table reports
an atomic mass for gallium of 69.72 amu.

In problem 11 we calculated
an atomic mass for gallium of 69.7231 amu,
which, when rounded to two decimal places, matches the value reported in the periodic table.

In problem 10 on the previous page, we saw that
silicon has three naturally occurring isotopes.

Isotope	Mass (amu)	Abundance (%)
Silicon-28	27.9769	92.23
Silicon-29	28.9765	4.67
Silicon-30	29.9738	3.10

The atomic mass of silicon is 28.0855 amu.

12. True or False?
In a sample of naturally occurring silicon, there is no individual silicon atom that has a mass of 28.0855 amu.

If true, explain why it's true.
If false, explain why its false.

Answer: True

In a sample of naturally occurring silicon,
92.23% of the individual atoms have mass of 27.9769 amu,
4.67% of the individual atoms have mass of 28.9765 amu, and
3.10% of the individual atoms have mass of 29.9738 amu.

As you can see, *none* of the individual atoms has a mass of 28.0855 amu.

Rather, the "atomic mass" of 28.0855 amu for silicon represents
a *weighted average* of the masses of the individual silicon atoms.

(This is why I mentioned earlier that it would be more accurate
to refer to the "atomic mass" as the "average atomic mass".)

**13. What concept do the decimal numbers in the periodic table represent?
What are the units for those decimal numbers?**

We have seen that the decimal numbers in the periodic table
represent the atomic mass for each element.

Therefore, the units for those decimal numbers are amu.

(In the next chapter on “Stoichiometry”,
we will learn an *alternative* way
to interpret the decimal numbers from the periodic table,
and alternative units that can be assigned to those numbers.)

Don’t confuse *atomic mass* with *mass number*.

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14. What is the definition of atomic mass?
What are the standard units for atomic mass?
What type of thing has an atomic mass?
Are atomic masses listed in the periodic table?

Naturally occurring elements are usually a *mixture* of different isotopes.

The *atomic mass* of an element is defined as
the *weighted average* of the masses of the individual isotopes in that mixture.

The standard units for atomic mass are amu.

Each *element* has its own atomic mass.
So, there is an atomic mass for hydrogen (1.008 amu), an atomic mass for helium (4.003 amu), an atomic mass for lithium (6.941 amu), an atomic mass for berylium (9.012 amu), etc.

The atomic mass for each element *is* listed in the periodic table.

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15. What is the definition of mass number?
What the units for mass number?
What type of thing has a mass number?
Are mass numbers listed in the periodic table?

As we learned in the lesson on Atomic Number and Mass Number,
mass number = number of protons + number of neutrons

Mass number is a unitless concept.

Every *isotope* has its own mass number.
So, a hydrogen atom with 1 proton and 0 neutrons has a mass number of 1 (so this isotope is called hydrogen-1);
a hydrogen atom with 1 proton and 1 neutron has a mass number of 2 (so this isotope is called hydrogen-2);
a hydrogen atom with 1 proton and 2 neutrons has a mass number of 3 (so this isotope is called hydrogen-3);
a helium atom with 2 protons and 1 neutron has a mass number of 3 (so this isotope is called helium-3);
a helium atom with 2 protons and 2 neutrons has a mass number of 4 (so this isotope is called helium-4);
etc.

Mass numbers are *not* listed in the periodic table.
(Mass numbers can not be listed in the periodic table, because the periodic table deals with elements, not with individual isotopes of the elements.)

We have seen that the mass of a proton is 1.007 amu.

The mass of a neutron is 1.009 amu.

The mass of an electron is 0.0005 amu,
roughly 2000 times smaller than the mass of a proton or neutron.

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16. What is the mass of a proton, rounded to the nearest whole number?
(Whole numbers are numbers like 0, 1, 2, 3, 4, 5, etc.)

What is the mass of a neutron, rounded to the nearest whole number?

What is the mass of an electron, rounded to the nearest whole number?

Rounded to the nearest whole number,
$1.007 \approx 1$
$1.009 \approx 1$
$0.0005 \approx 0$

So, rounded to the nearest whole numbers,
the mass of a proton is 1 amu,
the mass of a neutron is also 1 amu, and
the mass of an electron is 0 amu.

Notice that these approximations are quite accurate:
the mass of a proton or neutron, rounded to the first decimal place, is 1.0 amu;
and the mass of an electron, rounded to two decimal places, is 0.00 amu.

Therefore, if you want to calculate the mass of an isotope to the nearest whole number,
it is safe to use the following approximations:
mass of a proton ≈ 1 amu
mass of a neutron ≈ 1 amu
mass of an electron ≈ 0

What does it mean to approximate the mass of an electron as 0?
It means that we treat the mass of an electron as negligible, when compared with the mass of a proton or neutron.

We have seen that
naturally occurring magnesium is a *mixture* of
78.99% magnesium-24,
10.00% magnesium-25, and
11.01% magnesium-26.

One atom of magnesium-24 has a mass of 23.9850 amu.
One atom of magnesium-25 has a mass of 24.9858 amu.
One atom of magnesium-26 has a mass of 25.9826 amu.

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17. How many protons and neutrons does one atom of magnesium-24 contain?
What is the mass of one atom of magnesium-24, rounded to the nearest whole number?

How many protons and neutrons does one atom of magnesium-25 contain?
What is the mass of one atom of magnesium-25, rounded to the nearest whole number?

How many protons and neutrons does one atom of magnesium-26 contain?
What is the mass of one atom of magnesium-26, rounded to the nearest whole number?

From the periodic table, we know that any magnesium atom contains 12 protons.
24 is the mass number for magnesium-24.
mass number = number of protons + number of neutrons
So magnesium-24 contains 12 neutrons.
(For more practice with figuring out these numbers, review the previous lessons on Atomic Number and Mass Number and on Isotopes.)

We've seen that a proton has a mass of approximately 1 amu and a neutron also has a mass of approximately 1 amu; so, 12 protons and 12 neutrons result in a mass of approximately 24 amu; so magnesium-24 has a mass, rounded to the nearest whole number, of 24 amu.
(The mass of an electron is negligible, when compared to the mass of a proton or neutron, so we can disregard the mass of the electrons in this calculation.)
Alternatively, we were told above that one atom of magnesium-24 has a mass of 23.9850 amu; this confirms that, rounded to the nearest whole number, one atom of magnesium-24 has a mass of 24 amu.

From the periodic table, we know that any magnesium atom contains 12 protons.
25 is the mass number for magnesium-25.
mass number = number of protons + number of neutrons
So magnesium-25 contains 13 neutrons.

A proton has a mass of approximately 1 amu and a neutron also has a mass of approximately 1 amu; so, 12 protons and 13 neutrons result in a mass of approximately 25 amu; so magnesium-25 has a mass, rounded to the nearest whole number, of 25 amu.
Alternatively, we were told above that one atom of magnesium-25 has a mass of 24.9858 amu; this confirms that, rounded to the nearest whole number, one atom of magnesium-25 has a mass of 25 amu.

From the periodic table, we know that any magnesium atom contains 12 protons.
26 is the mass number for magnesium-26.
mass number = number of protons + number of neutrons
So magnesium-26 contains 14 neutrons.

A proton has a mass of approximately 1 amu and a neutron also has a mass of approximately 1 amu; so, 12 protons and 14 neutrons result in a mass of approximately 26 amu; so magnesium-26 has a mass, rounded to the nearest whole number, of 26 amu.
Alternatively, we were told above that one atom of magnesium-26 has a mass of 25.9826 amu; this confirms that, rounded to the nearest whole number, one atom of magnesium-26 has a mass of 26 amu.

Problem 17 gives us a new way to interpret the concept of "mass number".

In the earlier lesson on Atomic Number and Mass Number, we learned the definition
mass number = number of protons + number of neutrons

Now we can see an alternative, but equivalent, way to interpret the mass number of an isotope.
The mass number of an isotope can be interpreted as
the numerical value of the mass of one atom of the isotope, when measured in amu,
rounded to the nearest whole number.
That's why it's called the *mass* number!

But keep in mind that mass number is conventionally treated as a *unitless* concept.
"Mass" can take units of amu, mg, g, or kg.
But "mass number" is unitless.

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18. Consider the isotope fluorine-19.
Give three valid labels or interpretations for the number 19 in the name of this isotope.

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1. The number 19 represents the mass number for fluorine-19.

2. The number 19 represents the number of protons plus the number of neutrons in a single atom of fluorine-19.

3. The quantity 19 amu represents the mass of a single atom of fluorine-19, when rounded to the nearest whole number.

Let’s review the material we have discussed in this lesson.

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19. Define the term *mass*.

The *mass* of an object is defined as
the *quantity of matter* contained in that object.

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20. What are the units for mass that we have discussed in this lesson?
Arrange the units from biggest to smallest.
Give names and abbreviations.

Which units are most appropriate for ordinary sized objects,
and which units are most appropriate for atomic-scale objects.

We have discussed the following units for mass, arranged from biggest to smallest:
kilograms = kg
grams = g
milligrams = mg
atomic mass units = amu

Kilograms, grams, and milligrams are most convenient for ordinary-sized objects, and
amu are most appropriate for atomic-scale objects.

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21. True or False.
Naturally occurring elements are usually a *mixture* of different isotopes.

True.

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22. Define the term *atomic mass*.

Naturally occurring elements are usually a *mixture* of different isotopes.

The *atomic mass* of an element is defined as
the *weighted average* of the masses of the individual isotopes in that mixture.

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23. What are the standard units for atomic mass?

The standard units for atomic mass are amu.

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24. In the periodic table, what do the whole numbers and the decimal numbers represent?
Give both the *name* and the *meaning* of each concept.

The whole numbers in the periodic table represent the *atomic number* for each element.
atomic number = number of protons

The decimal numbers in the periodic table represent
the numerical value of the *atomic mass* for each element, expressed in amu.

The atomic mass for an element is the weighted average
of the masses of the individual isotopes in a naturally occurring sample of the element.

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25. Bromine has two naturally occurring isotopes, ^{79}Br (mass = 78.9183 amu, abundance = 50.69%) and ^{81}Br (mass = 80.9163 amu, abundance = 49.31%). Use this data to calculate the atomic mass of bromine.

Answer:
The atomic mass of bromine is 79.904 amu.

Solution:
? = atomic mass of bromine

We have the masses and natural abundances of each isotope:
Bromine-79: mass = 78.9183 amu, abundance = 50.69%
Bromine-81: mass = 80.9163 amu, abundance = 49.31%

Convert the percentages to decimals.

Now, we calculate the atomic mass,
which is a weighted average of the isotopic masses:
atomic mass = .5069(78.9183) + .4931(80.9163)
So, atomic mass = 79.904 amu

So, the atomic mass of bromine is 79.904 amu.

The abundances .5069 and .4931 are our weights.
Do these weights make sense?

Check:
Sum of weights = .5069 + .4931
So, sum of weights = 1
So, yes, the weights add up properly.

Does our answer for problem 25 make sense?

The atomic mass is a weighted average of 78.9183 amu and 80.9163 amu,
so we expect the result to be _between_ 78.9183 amu and 80.9163 amu.

Because the weights are nearly equal (about 51% and 49%),
we expect the result to be almost exactly halfway between those two values.

Our result for the weighted average was 79.904 amu,
which is indeed almost halfway between 78.9183 and 80.9163.
So, yes, our answer makes sense.

Hopefully you also checked your answer for problem 25
against the value in the periodic table.

You should find that the periodic table reports
an atomic mass for bromine of 79.90 amu.

In problem 25 we calculated
an atomic mass for bromine of 79.904 amu,
which, when rounded to two decimal places, matches the value reported in the periodic table.

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26. True or false?
Atomic mass is the same thing as mass number.

If the statement is true, explain why it's true.
If the statement is false, explain how atomic mass is different from mass number.

False.

mass number = number of neutrons + number of protons
Therefore, mass number is always a whole number.
Mass number is unitless.
Each isotope of an element has its own mass number.

The atomic mass of an element is a weighted average of the masses of all of the naturally occurring isotopes of that element.
Therefore, atomic mass is a decimal number.
The units for atomic mass are amu.
Each element has its own atomic mass; the atomic mass for each element is given by the decimal number in the periodic table.

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27. Consider the isotope sulfur-36.
Give three valid labels or interpretations for the number 36 in the name of this isotope.

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1. The number 36 represents the mass number for sulfur-36.

2. The number 36 represents the number of protons plus the number of neutrons in a single atom of sulfur-36.

3. The quantity 36 amu represents the mass of a single atom of sulfur-36, when rounded to the nearest whole number.

You’ve reached the end of the lesson.

You’re ready now to proceed to the next lesson for this chapter.

The next lesson is a Review Quiz, covering the material from the following lessons:
Atoms, Protons, Neutrons, and Electrons
Atomic Number and Mass Number
Isotopes
Atomic Mass