Chapter 2 Review Problems

INSTRUCTIONS:

You do not need to write the question, ONLY WRITE THE PROBLEM NUMBER and ANSWERS/SOLUTIONS.

- For problems that involve calculations, you must show your work to get full credit.
- For multiple choice questions, you can simply write the letter (a, b, c, or d) of the correct response.
- Use the navigation buttons at the bottom of the pages to get hints, check your answers, move to the next problem, or go back to previous pages.

Chapter Review Problems are due at the end of class period on the dates shown in the CHEM 108 Schedule.

- Late submissions will not be accepted unless the student can prove to the instructor that something outside of their control prevented them from turning in the problem set on the due date (see the course syllabus for more details).
2.1) Indicate whether each of the following statements about the nucleus of an atom is **true** or **false**.

a) The nucleus of an atom is neutral (total charge = 0).

b) The nucleus of an atom contains only neutrons.

c) The mass number is equal to the number of electrons present outside the nucleus.

d) The nucleus accounts for almost all the volume (size) of an atom.

e) The nucleus accounts for almost all the mass (weight) of an atom.

f) The nucleus can have a positive or negative charge depending on the identity of the atom.
2.1) Indicate whether each of the following statements about the nucleus of an atom is true or false.

a) The nucleus of an atom is neutral (total charge = 0).

b) The nucleus of an atom contains only neutrons.

c) The mass number is equal to the number of electrons present outside the nucleus.

d) The nucleus accounts for almost all the volume (size) of an atom.

e) The nucleus accounts for almost all the mass (weight) of an atom.

f) The nucleus can have a positive or negative charge depending on the identity of the atom.

HINT: Protons and neutrons are compacted together in the nucleus of atoms. Electrons are distributed in space around the nucleus and can be thought of as moving very fast in a volume surrounding the nucleus. The nucleus is the very small, dense core of an atom. An electron weighs about 1/2000th of a proton or a neutron.

Atoms are mostly empty space. Consider an analogy with familiar objects. If the nucleus (protons and neutrons) is represented by a golf ball placed on the pitcher’s mound of a baseball field, then the region occupied by electron(s) moving around the nucleus would be about the size of the baseball stadium and its parking lot!

An atom, as a whole, is neutral; it has no net (total) charge. The reason for this is that equal numbers of protons and electrons are always present in an atom.

For more help: See chapter 2 part 2 video or chapter 2 section 3 in the textbook.
2.1) Indicate whether each of the following statements about the nucleus of an atom is **true** or **false**.

a) The nucleus of an atom is neutral (total charge = 0).  **false**
   - The nucleus contain protons which are positive, and neutron, which are uncharged, therefore it has a net positive charge.

b) The nucleus of an atom contains only neutrons.  **false**
   - The nucleus of atom will always contain at least one proton.

c) The mass number is equal to the number of electrons present outside the nucleus.  **false**
   - The mass number is equal to the number of protons plus the number of neutrons.

d) The nucleus accounts for almost all the volume (size) of an atom.  **false**
   - The nucleus is the very small, dense core of an atom. It is the space that the electrons occupy that accounts for almost all the volume of an atom.

e) The nucleus accounts for almost all the mass (weight) of an atom.  **true**
   - The nucleus contain protons and neutrons; protons and neutrons are about 2000 times heavier than electrons.

f) The nucleus can have a positive or negative charge depending on the identity of the atom.  **false**
   - The nucleus contain protons which are positive, and neutron, which are uncharged, therefore it has a net positive charge.

For more details: See chapter 2 part 2 video or chapter 2 section 3 in the textbook.
2.2) Redraw the table below and then fill in the missing information for the particular atom in each of the column.

<table>
<thead>
<tr>
<th>Element</th>
<th>Mass number</th>
<th>Number of neutrons</th>
<th>Number of protons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Se</td>
<td>70</td>
<td>36</td>
<td>36</td>
</tr>
<tr>
<td></td>
<td>65</td>
<td>36</td>
<td>35</td>
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<td></td>
<td></td>
<td>36</td>
</tr>
</tbody>
</table>

**HINT:** The number of protons a particular atom contains determines that atom’s identity. We differentiate atoms with a particular number of protons by their names.

- For example, any atom that contains just one proton is called hydrogen. An atom with two protons is called helium. An atom with six protons is called carbon.

The mass number of an atom is defined as the number of protons plus the number of neutrons.

**For more help:** See chapter 2 part 2 video or chapter 2 section 3 in the textbook.
2.2) Redraw the table below and then fill in the missing information for the particular atom in each of the column.

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<td>34</td>
</tr>
<tr>
<td>Cu</td>
<td>65</td>
<td>36</td>
<td>29</td>
</tr>
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<td>Kr</td>
<td>71</td>
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**EXPLANATION:** The number of protons a particular atom contains determines that atom’s identity. We differentiate atoms with a particular number of protons by their names.

- For example, any atom that contains just one proton is called hydrogen. An atom with two protons is called helium. An atom with six protons is called carbon.

The mass number of an atom is defined as the number of protons plus the number of neutrons. Whenever you know the number of neutrons and the number of protons (atomic number or element name), you can determine the mass number.

\[
\text{mass number} = \text{number of protons} + \text{number of neutrons}
\]

Likewise, whenever you know the mass number and the number of protons (atomic number or element name), you can determine the number of neutrons. The reason for this is that rearranging the equation above to solve for the number of neutrons gives:

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

For more details: See chapter 2 part 2 video or chapter 2 section 3 in the textbook.
2.3) Which one of the following statements is true for an isotope of a particular element?

a) Isotopes of a particular element have the same number of neutrons but a different number of protons.

b) Isotopes of a particular element have the same number of protons but a different number of neutrons.

c) Isotopes of a particular element have a different number of protons and a different number of neutrons.
2.3) Which one of the following statements is true for an isotope of a particular element?

a) Isotopes of a particular element have the same number of neutrons but a different number of protons.

b) Isotopes of a particular element have the same number of protons but a different number of neutrons.

c) Isotopes of a particular element have a different number of protons and a different number of electrons.

HINT: See the chapter 2 part 2 video or read about isotopes in chapter 2 section 3 in the textbook.
2.3) Which one of the following statements is true for an isotope of a particular element?

- a) Isotopes of a particular element have the same number of neutrons but a different number of protons.

- b) Isotopes of a particular element have the same number of protons but a different number of neutrons. **TRUE**

- e) Isotopes of a particular element have a different number of protons and a different number of electrons.

**EXPLANATION:** Atoms of a particular element do not all have the same number of neutrons. Consider the element carbon as an example. Some carbon atoms have 6 neutrons, some have 7, and some have 8 neutrons. If you had a pile of carbon atoms in front of you and you were able to magically reach in and grab a single carbon atom, you would have about a 99% chance of grabbing a carbon with 6 neutrons, about a 1% chance of grabbing a carbon atom with 7 neutrons, and about a 0.1% chance of grabbing a carbon atom with 8 neutrons. These three different forms of carbon are called isotopes of carbon. Isotopes are defined as atoms with the same number of protons (same element), but a different number of neutrons.

**For more details:** See chapter 2 part 2 video or chapter 2 section 3 in the textbook.
2.4) An atom’s “atomic number” is the number of protons it contains and can be found on the periodic table. The mass number of an atom is defined as the number of protons plus the number of neutrons. When considering isotopes, a shorthand notation is often used in which the mass number is written as a superscript to the left of the atomic symbol. For example, consider the three isotopes of carbon in the table below.

<table>
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<tr>
<th>NUMBER OF NEUTRONS IN THE CARBON ATOM</th>
<th>SHorthand NOTATION</th>
</tr>
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<tbody>
<tr>
<td>6</td>
<td>(^{12}\text{C})</td>
</tr>
<tr>
<td>7</td>
<td>(^{13}\text{C})</td>
</tr>
<tr>
<td>8</td>
<td>(^{14}\text{C})</td>
</tr>
</tbody>
</table>

All carbon atoms have six protons (atomic number = 6). The number of neutrons vary, so the mass number varies.

**QUESTION:** \(^{18}\text{F}\) is administered to patients in order to monitor brain activity. How many neutrons are in \(^{18}\text{F}\)?
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All carbon atoms have six protons (atomic number = 6). The number of neutrons vary, so the mass number varies.

QUESTION: $^{18}\text{F}$ is administered to patients in order to monitor brain activity. How many neutrons are in $^{18}\text{F}$?

**HINT:** Whenever you know the number of neutrons and the number of protons (atomic number or element name), you can determine the mass number.

\[
\text{mass number} = \text{number of protons} + \text{number of neutrons}
\]

Likewise, whenever you know the mass number and the number of protons (atomic number or element name), you can determine the number of neutrons. The reason for this is that rearranging the equation above to solve for the number of neutrons gives:

\[
\text{number of neutrons} = \text{mass number} - \text{number of protons}
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2.4) An atom’s “atomic number” is the number of protons it contains and can be found on the periodic table. The mass number of an atom is defined as the number of protons plus the number of neutrons. When considering isotopes, a shorthand notation is often used in which the mass number is written as a superscript to the left of the atomic symbol. For example, consider the three isotopes of carbon in the table below.

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<td>$^{14}$C</td>
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</table>

All carbon atoms have six protons (atomic number = 6). The number of neutrons vary, so the mass number varies.

QUESTION: $^{18}$F is administered to patients in order to monitor brain activity. How many neutrons are in $^{18}$F?

ANSWER: 9 neutrons

EXPLANATION: mass number = number of protons + number of neutrons

All fluorine atoms have 9 protons (atomic number = 9). So for $^{18}$F:

$$18 = 9 + \text{number of neutrons}$$

Rearranging the equation above to solve for the number of neutrons gives:

$$\text{number of neutrons} = 18 - 9 = 9 \text{ neutrons}$$

For more details: See chapter 2 part 2 video or chapter 2 section 3 in the textbook.
2.5) Classify each of these elements as metal or nonmetal.

a) Ne
b) F
c) Na
d) Fe
e) Cr
2.5) Classify each of these elements as metal or nonmetal.

a) Ne  
b) F  
c) Na  
d) Fe  
e) Cr

**HINT:** Use the element’s *position in the periodic table* to determine if it is a metal or nonmetal.

**For more help:** See chapter 2 part 5 video or chapter 2 section 5 in the textbook.
2.5) Classify each of these elements as metal or nonmetal.

a) Ne: nonmetal
b) F: nonmetal
c) Na: metal
d) Fe: metal
e) Cr: metal

EXPLANATION: Use the element’s position in the periodic table to determine if it is a metal or nonmetal.
2.6) A sample contains $1.43 \times 10^{24}$ carbon atoms. How many moles of carbon are in the sample?

REMINDER: Be sure to use the correct number of significant figures in your answer.
2.6) A sample contains $1.43 \times 10^{24}$ carbon atoms. How many moles of carbon are in the sample?

**HINT:** This is a *unit conversion problem*. Convert from units of “atoms” to units of “moles.” You need to know the relationship between atoms and moles; we use Avogadro’s Number: $6.022 \times 10^{23}$ atoms $= 1$ mole.

**For more help:** See chapter 2 part 3 video or chapter 2 section 4 in the textbook.
2.6) A sample contains $1.43 \times 10^{24}$ carbon atoms. How many moles of carbon are in the sample?

ANSWER: 2.37 moles

Your answer should have three significant figures.
2.6) A sample contains $1.43 \times 10^{24}$ carbon atoms. How many moles of carbon are in the sample?

**ANSWER:** 2.37 moles

---

**EXPLANATION:** The relationship between atoms and moles, $6.022 \times 10^{23}$ atoms = 1 mole, is used as a conversion factor.

\[
\frac{1.43 \times 10^{24} \text{ C atoms}}{6.022 \times 10^{23} \text{ C atoms}} = 2.37 \text{ moles C}
\]

**For more details:** See chapter 2 part 3 video or chapter 2 section 4 in the textbook.
2.7) A sample contains $1.65 \times 10^{24}$ zinc atoms. How many moles of zinc are in the sample?
2.7) A sample contains $1.65 \times 10^{24}$ zinc atoms. How many moles of zinc are in the sample?

**HINT:** This is a *unit conversion problem*. Convert from units of “atoms” to units of “moles.” You need to know the relationship between atoms and moles; we use Avogadro’s Number: $6.022 \times 10^{23}$ atoms = 1 mole.

**For more help:** See chapter 2 part 3 video or chapter 2 section 4 in the textbook.
2.7) A sample contains $1.65 \times 10^{24}$ zinc atoms. How many moles of zinc are in the sample?

**ANSWER:** 2.74 moles

Your answer should have three significant figures.

[CLICK HERE to see the complete solution for this problem](#)
2.7) A sample contains $1.65 \times 10^{24}$ zinc atoms. How many moles of zinc are in the sample?

**ANSWER:** 2.74 moles

**EXPLANATION:** The relationship between atoms and moles, $6.022 \times 10^{23}$ atoms = 1 mole, is used as a conversion factor.

For more details: See chapter 2 part 3 video or chapter 2 section 4 in the textbook.
2.8) A sample contains 0.75 moles of silicon (Si). How many silicon *atoms* are in the sample?
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**HINT:** This is a unit conversion problem. Convert from units of “moles” to units of “atoms.” You need to know the relationship between moles and atoms; we use Avogadro’s Number: \(6.022 \times 10^{23}\) atoms = 1 mole.

**For more help:** See chapter 2 part 3 video or chapter 2 section 4 in the textbook.
2.8) A sample contains 0.75 moles of silicon (Si). How many silicon \textit{atoms} are in the sample?

ANSWER: \(4.5 \times 10^{23}\) Si atoms

Your answer should have \textit{two} significant figures.
2.8) A sample contains 0.75 moles of silicon (Si). How many silicon atoms are in the sample?

**ANSWER:** \(4.5 \times 10^{23}\) Si atoms

**EXPLANATION:** The relationship between atoms and moles, \(6.022 \times 10^{23}\) atoms = 1 mole, is used as a *conversion factor*.

\[
\frac{0.75 \text{ moles Si}}{1 \text{ mole Si}} \times \frac{6.022 \times 10^{23} \text{ Si atoms}}{1 \text{ mole Si}} = 4.5 \times 10^{23} \text{ Si atoms}
\]

**For more details:** See chapter 2 part 3 video or chapter 2 section 4 in the textbook.
2.9) A sample contains 0.33 moles of Magnesium (Mg). How many magnesium atoms are in the sample?
2.9) A sample contains 0.33 **moles** of Magnesium (Mg). How many magnesium **atoms** are in the sample?

**HINT:** This is a *unit conversion problem*. Convert from units of “**moles**” to units of “**atoms**.” You need to know the relationship between moles and atoms; we use Avogadro’s Number: \(6.022 \times 10^{23} \text{ atoms} = 1 \text{ mole}\).

**For more help:** See chapter 2 part 3 video or chapter 2 section 4 in the textbook.
2.9) A sample contains 0.33 moles of Magnesium (Mg). How many magnesium atoms are in the sample?

Your answer should have two significant figures.

ANSWER: $2.0 \times 10^{23}$ Mg atoms

CLICK HERE to see the complete solution for this problem
2.9) A sample contains 0.33 moles of Magnesium (Mg). How many magnesium atoms are in the sample?

**ANSWER:** $2.0 \times 10^{23}$ Mg atoms

**EXPLANATION:** The relationship between atoms and moles, $6.022 \times 10^{23}$ atoms = 1 mole, is used as a conversion factor.

For more details: See chapter 2 part 3 video or chapter 2 section 4 in the textbook.
2.10) A sample contains 1.78 moles of magnesium (Mg). How many grams of magnesium are in the sample?

**IMPORTANT NOTE:** In order for your answers to exactly match the solutions provided for review problems (and worksheet problems), round molar masses to two digits to the right of the decimal point when possible.

- For example, use 24.31 grams/mole for the molar mass of magnesium.
HINT: This is a unit conversion problem. Convert from units of “moles” to units of “grams.” The relationship between moles and grams of a particular element is called the molar mass. The molar mass for each element is written under the element’s name/symbol in the periodic table.

For more help: See chapter 2 part 4 video or chapter 2 section 4 in the textbook.

2.10) A sample contains 1.78 moles of magnesium (Mg). How many grams of magnesium are in the sample?
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ANSWER: 43.3 grams Mg

Your answer should have three significant figures.
2.10) A sample contains 1.78 moles of magnesium (Mg). How many grams of magnesium are in the sample?

**ANSWER:** 43.3 grams Mg

**EXPLANATION:** The relationship between moles and grams is the molar mass. The molar mass of magnesium is 24.31 g/mole (1 mole Mg = 24.31 grams Mg); this quantity is used as a conversion factor.

For more details: See chapter 2 part 4 video or chapter 2 section 4 in the textbook.
2.11) A sample contains 5.18 moles of gold (Au). How many grams of gold are in the sample?

REMINDER: Round molar masses to two digits to the right of the decimal point when possible.
2.11) A sample contains 5.18 moles of gold (Au). How many grams of gold are in the sample?

**HINT:** This is a unit conversion problem. Convert from units of “moles” to units of “grams.” The relationship between moles and grams of a particular element is called the molar mass. The molar mass for each element is written under the element’s name/symbol in the periodic table.

**For more help:** See chapter 2 part 4 video or chapter 2 section 4 in the textbook.
2.11) A sample contains 5.18 moles of gold (Au). How many grams of gold are in the sample?

ANSWER: 1020 grams Au

Your answer should have three significant figures. The absence of a decimal point in “1020” indicates that the right-most zero is not significant.
2.11) A sample contains 5.18 moles of gold (Au). How many grams of gold are in the sample?

**ANSWER: 1020 grams Au**

**EXPLANATION:** The relationship between moles and grams is the molar mass. The molar mass of gold is 196.97 g/mole (1 mole Au = 196.97 grams Au); this quantity is used as a conversion factor.

For more details: See chapter 2 part 4 video or chapter 2 section 4 in the textbook.
2.12) A sample contains 0.436 grams of helium (He). How many moles of helium are in the sample?
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**HINT:** This is a unit conversion problem. Convert from units of “grams” to units of “moles.” The relationship between grams and moles of a particular element is called the molar mass. The molar mass for each element is written under the element’s name/symbol in the periodic table.

For more help: See chapter 2 part 4 video or chapter 2 section 4 in the textbook.
2.12) A sample contains 0.436 \textit{grams} of helium (He). How many \textit{moles} of helium are in the sample?

\textbf{ANSWER:} 0.109 moles He

Your answer should have \textit{three} significant figures.
2.12) A sample contains 0.436 grams of helium (He). How many moles of helium are in the sample?

**ANSWER:** 0.109 moles He

**EXPLANATION:** The relationship between moles and grams is the molar mass. The molar mass of helium is 4.00 g/mole (1 mole He = 4.00 grams He); this quantity is used as a conversion factor.

\[
\frac{0.436 \text{ grams He}}{4.00 \text{ grams He/mole}} = 0.109 \text{ moles He}
\]

**For more details:** See chapter 2 part 4 video or chapter 2 section 4 in the textbook.
2.13) How many calcium \textit{atoms} are present in 129 \textit{grams} of calcium (Ca)?
2.13) How many calcium atoms are present in 129 grams of calcium (Ca)?

**HINT:** This is a TWO-STEP unit conversion problem. First, convert from grams to moles, and then convert from moles to atoms. You may find the conversion map shown below to be helpful when doing atoms/moles/grams calculations:

**REMINDER:** Round molar masses to two digits to the right of the decimal point when possible.

**For more help:** See chapter 2 part 4 video or chapter 2 section 4 in the textbook.
2.13) How many calcium *atoms* are present in 129 *grams* of calcium (Ca)?

**ANSWER:** \(1.94 \times 10^{24}\) Ca atoms

Your answer should have *three* significant figures.
2.13) How many calcium atoms are present in 129 grams of calcium (Ca)?

ANSWER: \(1.94 \times 10^{24}\) Ca atoms

EXPLANATION: This is a TWO STEP unit conversion problem.

First step: convert from grams to moles:

\[
\begin{array}{c|c|c}
\text{129 grams Ca} & \text{1 mole Ca} & \text{3.218562 moles Ca} \\
\text{40.08 grams Ca} & & \text{(unrounded)}
\end{array}
\]

NOTE: When doing two-step conversion, do not round the calculator’s values until the second step.

Second step: convert from moles to atoms:

\[
\begin{array}{c|c|c}
\text{3.218562 moles Ca} & \text{6.022 \times 10^{23} Ca atoms} & \text{1.94 \times 10^{24} Ca atoms} \\
\text{1 mole Ca} & & 
\end{array}
\]

Alternatively, these two steps can be combined into one equation:

\[
\begin{array}{c|c|c}
\text{129 grams Ca} & \text{1 mole Ca} & \text{6.022 \times 10^{23} Ca atoms} \\
\text{40.08 grams Ca} & \text{1 mole Ca} & \text{1.94 \times 10^{24} Ca atoms}
\end{array}
\]

For more details: See chapter 2 part 4 video or chapter 2 section 4 in the textbook.
2.14) What is the mass (grams) of $2.5 \times 10^{19}$ silver (Ag) atoms?
2.14) What is the mass (grams) of $2.5 \times 10^{19}$ silver (Ag) atoms?

**HINT:** This is a **TWO-STEP unit conversion problem**. First, convert from **atoms** to **moles**, and then convert from **moles** to **grams**. You may find the conversion map shown below to be helpful when doing atoms/moles/grams calculations:

REMINDER: Round molar masses to **two digits** to the right of the decimal point when possible.

**For more help:** See chapter 2 part 4 video or chapter 2 section 4 in the textbook.
2.14) What is the mass (grams) of $2.5 \times 10^{19}$ silver (Ag) atoms?

**ANSWER:** $4.5 \times 10^{-3}$ grams Ag (or 0.0045 grams Ag)

Your answer should have *two* significant figures.

CLICK HERE to see the complete solution for this problem.
2.14) What is the mass (grams) of $2.5 \times 10^{19}$ silver (Ag) atoms?  

**ANSWER:** $4.5 \times 10^{-3}$ grams Ag (or 0.0045 grams Ag)

**EXPLANATION:** This is a **TWO STEP** unit conversion problem.

**First step:** convert from *atoms* to *moles*:

\[
\begin{array}{c|c}
\text{2.5} \times 10^{19} \text{ Ag atoms} & 1 \text{ mole Ag} \\
6.022 \times 10^{23} \text{ Ag atoms} &
\end{array}
\]

\[= 4.1514447 \times 10^{-5} \text{ moles Ag (unrounded)}\]

**NOTE:** When doing two-step conversion, do not round the calculator’s values until the second step.

**Second step:** convert from *moles* to *grams*:

\[
\begin{array}{c|c}
4.1514447 \times 10^{-5} \text{ moles Ag} & 107.87 \text{ grams Ag} \\
1 \text{ mole Ag} &
\end{array}
\]

\[= 4.5 \times 10^{-3} \text{ grams Ag (or 0.0045 grams Ag)}\]

**Alternatively,** these two steps can be combined into one equation:

\[
\begin{array}{c|c|c}
\text{2.5} \times 10^{19} \text{ Ag atoms} & 1 \text{ mole Ag} & 107.87 \text{ grams Ag} \\
6.022 \times 10^{23} \text{ Ag atoms} & 1 \text{ mole Ag} &
\end{array}
\]

\[= 4.5 \times 10^{-3} \text{ grams Ag}\]

For more details: See chapter 2 part 4 video or chapter 2 section 4 in the textbook.
2.15) How many sulfur \textit{atoms} are present in 151 \textit{grams} of sulfur (S)?
2.15) How many sulfur *atoms* are present in 151 *grams* of sulfur (S)?

**HINT:** This is a **TWO-STEP unit conversion problem**. First, convert from *grams* to *moles*, and then convert from *moles* to *atoms*. You may find the conversion map shown below to be helpful when doing atoms/moles/grams calculations:

![Conversion map diagram]

**REMINDER:** Round molar masses to **two digits** to the right of the decimal point when possible.

**For more help:** See chapter 2 part 4 video or chapter 2 section 4 in the textbook.
2.15) How many sulfur \textit{atoms} are present in 151 \textit{grams} of sulfur (S)?

\textbf{ANSWER: } \textit{2.84} \times 10^{24} \text{ S atoms}

Your answer should have \textit{three} significant figures.
2.15) How many sulfur atoms are present in 151 grams of sulfur (S)?

EXPLANATION: This is a TWO STEP unit conversion problem.

**First step:** convert from grams to moles:

\[
\begin{array}{c|c}
151 \text{ grams S} & 1 \text{ mole S} \\
32.07 \text{ grams S} & \\
\end{array}
\]

\[
= \frac{4.7088450 \text{ moles S}}{} \quad \text{(unrounded)}
\]

NOTE: When doing two-step conversion, do not round the calculator’s values until the second step.

**Second step:** convert from moles to atoms:

\[
\begin{array}{c|c}
4.7088450 \text{ moles S} & 6.022 \times 10^{23} \text{ S atoms} \\
1 \text{ mole S} & \\
\end{array}
\]

\[
= \frac{2.84 \times 10^{24} \text{ S atoms}}{}
\]

Alternatively, these two steps can be combined into one equation:

\[
\begin{array}{c|c|c}
151 \text{ grams S} & 1 \text{ mole S} & 6.022 \times 10^{23} \text{ S atoms} \\
32.07 \text{ grams S} & 1 \text{ mole S} & \\
\end{array}
\]

\[
= \frac{2.84 \times 10^{24} \text{ S atoms}}{}
\]

**ANSWER:** \(2.84 \times 10^{24} \text{ S atoms}\)

For more details: See chapter 2 part 4 video or chapter 2 section 4 in the textbook.
2.16) What is the mass (grams) of $5.77 \times 10^{23}$ argon (Ar) atoms?
2.16) What is the mass (grams) of 5.77 \times 10^{23} \text{ argon (Ar) atoms?}

**HINT:** This is a **TWO-STEP unit conversion problem**. First, convert from **atoms** to **moles**, and then convert from **moles** to **grams**. You may find the conversion map shown below to be helpful when doing atoms/moles/grams calculations:

**REMINDER:** Round **molar masses** to **two digits** to the right of the decimal point when possible.

For more help: See chapter 2 part 4 video or chapter 2 section 4 in the textbook.
2.16) What is the mass (grams) of $5.77 \times 10^{23}$ argon (Ar) atoms?

ANSWER: 38.3 grams

Your answer should have three significant figures.

CLICK HERE to see the complete solution for this problem.
2.16) What is the mass (grams) of $5.77 \times 10^{23}$ argon (Ar) atoms?

**ANSWER:** 38.3 grams

**EXPLANATION:** This is a **TWO STEP** unit conversion problem.

**First step:** convert from *atoms* to *moles*:

\[
\frac{5.77 \times 10^{23} \text{ Ar atoms}}{6.022 \times 10^{23} \text{ Ar atoms}} = 0.9581534 \text{ moles Ar (unrounded)}
\]

**NOTE:** When doing two-step conversion, do not round the calculator’s values until the second step.

**Second step:** convert from *moles* to *grams*:

\[
\frac{0.9581534 \text{ moles Ar}}{1 \text{ mole Ar}} = 38.3 \text{ grams Ar}
\]

**Alternatively**, these two steps can be combined into one equation:

\[
\frac{5.77 \times 10^{23} \text{ Ar atoms}}{6.022 \times 10^{23} \text{ atoms}} \times 1 \text{ mole Ar} \times 39.95 \text{ grams Ar} = 38.3 \text{ grams Ar}
\]

**For more details:** See chapter 2 part 4 video or chapter 2 section 4 in the textbook.