Chapter 3 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).
3.1) Classify the atomic orbitals as either s, p, or d.

a) 

b) 

c) 

Click here for a hint

Click here to check your answer

Go to next question
3.1) Classify the atomic orbitals as either $s$, $p$, or $d$.

HINT: Atomic orbitals are classified as $s$, $p$, or $d$ based on their shape.

For more help: See chapter 3 part 2 video or chapter 3 section 2 in the textbook.
3.1) Classify the atomic orbitals as either \( s \), \( p \), or \( d \).

a) \( p \) orbitals

b) \( s \) orbital

c) \( d \) orbitals

EXPLANATION:
Atomic orbitals are classified as \( s \), \( p \), or \( d \) based on their shape.

For more details: See chapter 3 part 2 video or chapter 3 section 2 in the textbook.
3.2) Draw an “energy level diagram” for a nitrogen atom.
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HINT: A nitrogen atom 7 electrons. An empty (no electrons) energy level diagram for multi-electron atoms is shown on the left. Electrons are arranged (configured) into the orbitals in the way that results in the lowest possible energy. Nature does this by obeying the following principles:

1) The Aufbau Principle:” "Aufbau" (German) means build-up or construct. The aufbau principle states that an electron occupies the lowest energy orbital that can receive it.

2) The Pauli Exclusion Principle: An orbital can hold a maximum of two electrons. When two electrons occupy the same orbital, one electron has spin “up” the other has spin “down.” Having two electrons in the same orbital with opposite spin states is lower in energy than when both spins are the same.

3) Hund’s Rules: (i) When electrons are configured into orbitals that all have the same energy, for example the \(2p_x, 2p_y, \) and \(2p_z\), a single electron is placed into each of the equal-energy orbitals before a second electron is added to an occupied orbital. (ii) When electrons are configured into a set of orbitals that all have the same energy, the spins of the first electrons to be placed into each orbital are all in the same state (for example all “up”).

For more help: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.
3.2) Draw an “energy level diagram” for a nitrogen atom.

**EXPLANATION:** A nitrogen atom has 7 protons and therefore 7 electrons. We use the multi-electron energy level diagram.

- **The first and second electrons** occupy the 1s orbital according to the **Aufbau Principle**.
  - Note that the nitrogen energy level diagram does not include many high energy orbitals (n > 2) because nitrogen **only has seven electrons**.

- Because the 1s orbital is now full (orbitals hold a maximum of two electrons), **the third and fourth electrons** occupy the 2s orbital.

- **The fifth electron** goes into any one of the three 2p orbitals. The three 2p orbitals all are equivalent in energy. We will arbitrarily choose to put the **fifth electron** into the 2px orbital (represented by the red arrow).

- The **sixth and seventh electrons** (represented by the blue arrows) each go into **unoccupied** 2p orbitals. Recall **Hund’s Rule (i)**: when electrons are configured into orbitals that all have the same energy, a **single electron** is placed into each of the equal-energy orbitals **before** the electrons are added to an occupied orbital.

- How did we know to put the spin state of the sixth and seventh electrons in the “up” state? **Hund’s Rule (ii)** explains that when electrons are configured into a set of orbitals that have the same energy, the spins of the first electrons to be placed into each orbital are all in the same state (for example all “up”).

**For more details:** See chapter 3 part 3 video or chapter 3 section 2 in the textbook.
3.3) Draw an “energy level diagram” for a calcium (Ca) atom.
HINT: Use the periodic table to determine the number of electrons that are present in a krypton atom. An empty (no electrons) energy level diagram for multi-electron atoms is shown on the left. Electrons are arranged (configured) into the orbitals in the way that results in the lowest possible energy. Nature does this by obeying the Aufbau Principle, the Pauli Exclusion Principle, and Hund’s Rules. See the hint from the previous question (3.2) for more details.

For more help: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.
EXPLANATION: A calcium atom has 20 protons and therefore 20 electrons. Electrons are arranged (configured) into the orbitals in the way that results in the lowest possible energy. Nature does this by obeying the Aufbau Principle, the Pauli Exclusion Principle, and Hund’s Rules. See the hint from the previous question (3.2) for more details. Note that the 4s orbital fills before the 3d orbital because the 4s orbital lower in energy than the 3d orbitals (the 3d orbitals are not shown here because they do not contain electrons).

For more details: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.
3.4) In a situation where an electron goes from the $n = 1$ orbital (ground state) to an $n = 3$ orbital, which of the following statements are true?

- a) Energy was released by the atom.
- b) Energy was absorbed by the atom.
- c) There was no change in the atom’s energy.
3.4) In a situation where an electron goes from the \( n = 1 \) orbital (ground state) to an \( n = 3 \) orbital, which of the following statements are true?

- a) Energy was released by the atom.
- b) Energy was absorbed by the atom.
- c) There was no change in the atom’s energy.

**HINT:** Look at the energy level diagram. Base your answer on the relative energy of the \( n=1 \) orbital (1s) to any of the \( n = 3 \) orbitals (3s, 3p, or 3d).
3.4) In a situation where an electron goes from the n = 1 orbital (ground state) to an n = 3 orbital, which of the following statements are true?

a) Energy was released by the atom.  **false**

b) Energy was absorbed by the atom.  **true**

c) There was no change in the atom’s energy.  **false**

EXPLANATION:
An atom must **absorb** a quantity of energy indicated by one of the blue arrows in order to move from the n=1 orbital to an n = 3 orbital.
3.5) Label each of these orbital designations as “possible” (exists) or “impossible” (does not exist).

a) 1s
b) 2s
c) 1p
d) 2p
e) 2d
f) 3d
3.5) Label each of these orbital designations as "possible" (exists) or "impossible" (does not exist).

a) 1s
b) 2s
c) 1p
d) 2p
e) 2d
f) 3d

HINT:
Check the energy level diagram to see if these orbitals are present (possible) or if they do not exist (impossible).

For more help: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.
3.5) Label each of these orbital designations as “possible” (exists) or “impossible” (does not exist).

a) 1s  possible
b) 2s  possible
c) 1p  impossible
   1p orbitals do not exist
d) 2p  possible
e) 2d  impossible
   2d orbitals do not exist
f) 3d  possible

For more details: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.
3.6) Redraw the table below and then enter the number of electrons in each quantum level \((n)\) for the following elements. If the quantum level does not contain any electrons, enter a “0” (zero).

<table>
<thead>
<tr>
<th>Element</th>
<th>(n = 1)</th>
<th>(n = 2)</th>
<th>(n = 3)</th>
<th>(n = 4)</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen (H)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>magnesium (Mg)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>bromine (Br)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>potassium (K)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
3.6) Redraw the table below and then enter the number of electrons in each quantum level \((n)\) for the following elements. If the quantum level does not contain any electrons, enter a “0” (zero).

**HINT:** For each element in the table, arrange the electrons into the energy level diagram. Next, count the number of electrons in each quantum level \((n)\).

<table>
<thead>
<tr>
<th>Element</th>
<th>(n = 1)</th>
<th>(n = 2)</th>
<th>(n = 3)</th>
<th>(n = 4)</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen ((\text{H}))</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>magnesium ((\text{Mg}))</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>bromine ((\text{Br}))</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>potassium ((\text{K}))</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

For more help: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.
3.6) Redraw the table below and then enter the number of electrons in each quantum level (n) for the following elements. If the quantum level does not contain any electrons, enter a “0” (zero).

<table>
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<th>Element</th>
<th>n = 1</th>
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<th>n = 3</th>
<th>n = 4</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen (H)</td>
<td>1</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>magnesium (Mg)</td>
<td>2</td>
<td>8</td>
<td>2</td>
<td>0</td>
</tr>
<tr>
<td>bromine (Br)</td>
<td>2</td>
<td>8</td>
<td>18</td>
<td>7</td>
</tr>
<tr>
<td>potassium (K)</td>
<td>2</td>
<td>8</td>
<td>8</td>
<td>1</td>
</tr>
</tbody>
</table>

Note that the 4s orbital is lower in energy than the 3d orbitals, therefore the 4s orbital will accept one electron before electrons are placed in the 3d orbitals.

EXPLANATION: For each element in the table, electrons are placed into the energy level diagram. After doing so, you can count the number of electrons in each quantum level (n).
3.7) Determine the number of *valence electrons* in each of the atoms listed below.

a) a hydrogen (H) atom  
b) a carbon (C) atom  
c) an argon (Ar) atom  
d) a sulfur (S) atom  
e) a barium (Ba) atom  
f) a chlorine (Cl) atom
3.7) Determine the number of **valence electrons** in each of the atoms listed below.

a) a hydrogen (H) atom  
b) a carbon (C) atom  
c) an argon (Ar) atom  
d) a sulfur (S) atom  
e) a barium (Ba) atom  
f) a chlorine (Cl) atom

**HINT:** For **s-block** and **p-block** elements, the number of valence electrons can be determined by the element’s **position in the periodic table**.

For more help: See chapter 3 part 4 video or chapter 3 section 3 in the textbook.
3.7) Determine the number of valence electrons in each of the atoms listed below.

EXPLANATION: For s-block and p-block elements, the number of valence electrons is the same as the element’s column group number. The group numbers are shown as roman numerals above the column of elements.

a) a hydrogen (H) atom 1
b) a carbon (C) atom 4
c) an argon (Ar) atom 8
d) a sulfur (S) atom 6
e) a barium (Ba) atom 2
f) a chlorine (Cl) atom 7

For more details: See chapter 3 part 4 video or chapter 3 section 3 in the textbook.
3.8) Determine whether the following atoms will **gain** or **lose** electron when forming an ion.

a) an oxygen (O) atom  
b) a calcium (Ca) atom  
c) a fluorine (F) atom  
d) a sodium (Na) atom  
e) a nitrogen (N) atom
3.8) Determine whether the following atoms will \textbf{gain} or \textbf{lose} electron when forming an ion.

a) an oxygen (O) atom
b) a calcium (Ca) atom
c) a fluorine (F) atom
d) a sodium (Na) atom
e) a nitrogen (N) atom

\textbf{HINT:}

Metal atoms \textit{lose} electrons to form \textit{positive ions}.
Nonmetal atoms \textit{gain} electrons to form \textit{negative ions}.

\textbf{For more help:} See \textit{chapter 3 part 5 video} or chapter 3 section 4 in the textbook.
3.8) Determine whether the following atoms will **gain** or **lose** electron when forming an ion.

**EXPLANATION:**

Metal atoms **lose** electrons to form **positive ions**. If an atom **loses** one or more electrons, it will then have more protons than electrons and have an overall **positive charge**. **Positive ions** are called **cations**.

Nonmetal atoms **gain** electrons to form **negative ions**. If an atom **gains** one or more electrons, it will then have more electrons than protons and have an overall **negative charge**. **Negative ions** are called **anions**.

For more details: See chapter 3 part 5 video or chapter 3 section 4 in the textbook.
3.9) Predict the charge of the following ions.

a) a sodium ion
b) an oxide ion
c) a calcium ion
d) a fluoride ion
e) a potassium ion
f) a nitride ion
3.9) Predict the charge of the following ions.

a) a sodium ion
b) an oxide ion
c) a calcium ion
d) a fluoride ion
e) a potassium ion
f) a nitride ion

**HINT:** Ions generally form such that the *ion* has an **octet of electrons** in its outermost shell. This tendency will allow us to predict the charge of the ion that is formed for particular elements.

**For more help:** See [chapter 3 part 5 video](#) or chapter 3 section 4 in the textbook.

### Charge for s-Block Cations and p-Block Nonmetal Anions

<table>
<thead>
<tr>
<th>Periodic Group</th>
<th>Number of Valence Electrons of the Element</th>
<th>Number of Electrons Gained or Lost in Ion Formation</th>
<th>Charge of Ion Formed</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>s-Block Elements</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Group I</td>
<td>1</td>
<td>Lose 1 electron</td>
<td>1+</td>
</tr>
<tr>
<td>Group II</td>
<td>2</td>
<td>Lose 2 electrons</td>
<td>2+</td>
</tr>
<tr>
<td><strong>p-Block Nonmetal Elements</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Group III</td>
<td>There are no Group III non-metals (only metals and metalloids)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Group IV</td>
<td>4</td>
<td>Do not form ions, high energy to gain or lose 4 electrons!</td>
<td></td>
</tr>
<tr>
<td>Group V</td>
<td>5</td>
<td>Gain 3 electrons</td>
<td>3-</td>
</tr>
<tr>
<td>Group VI</td>
<td>6</td>
<td>Gain 2 electrons</td>
<td>2-</td>
</tr>
<tr>
<td>Group VII</td>
<td>7</td>
<td>Gain 1 electron</td>
<td>1-</td>
</tr>
<tr>
<td>Group VIII</td>
<td>8</td>
<td>Do not form ions, noble gas atoms have filled outer shells.</td>
<td></td>
</tr>
</tbody>
</table>
3.9) Predict the charge of the following ions.

a) a sodium ion  \(1^+\)
b) an oxide ion  \(2^-\)
c) a calcium ion  \(2^+\)
d) a fluoride ion  \(1^-\)
e) a potassium ion  \(1^+\)
f) a nitride ion  \(3^-\)

**EXPLANATION:**
Ions generally form such that the ion has an octet of electrons in its outermost shell. This tendency will allow us to predict the charge of the ion that is formed for particular elements.
3.10) Redraw the table below and then fill in the missing information for the particular species 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>
<th>Number of electrons</th>
<th>charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>Se</td>
<td>76</td>
<td>36</td>
<td>36</td>
<td>36</td>
<td>2-</td>
</tr>
<tr>
<td></td>
<td>65</td>
<td></td>
<td>39</td>
<td></td>
<td>1+</td>
</tr>
</tbody>
</table>

**NOTE:** You did a similar problem in chapter 2. In that problem, all of the species were atoms. All atoms are electrically neutral (charge = 0) because the number of protons is the same as the number of electrons. In this problem, one or more of the species are ions.
3.10) Redraw the table below and then fill in the missing information for the particular species in each of the column.

<table>
<thead>
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</tr>
<tr>
<td></td>
<td></td>
<td>39</td>
<td></td>
<td>36</td>
<td></td>
</tr>
</tbody>
</table>

HINT: You did a similar problem in chapter 2. In that problem, all of the species were atoms. All atoms are electrically neutral (total charge = 0) because the number of protons is the same as the number of electrons. In this problem, one or more of the species are ions. Cations have more protons than electrons; this gives them a positive charge. Anions have more electrons than protons; this gives them a negative charge.

The mass number of an atom is defined as the number of protons plus the number of neutrons. Whenever you know the two of the three quantities in the equations shown below, you can determine the value for the third quantity.

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

For more help: See chapter 3 part 5 video or chapter 3 section 4 in the textbook.
EXPLANATION: All atoms are electrically neutral (total charge = 0) because the number of protons is the same as the number of electrons. In this problem, Kr is an atom because it has the same number of protons and electrons.

Cations have more protons than electrons; this gives them a net positive charge. Anions have more electrons than protons; this gives them a net negative charge. In this problem, Se and Cu are ions. You were told that Se has a 2- charge, it will therefore have two more electrons than it does protons. You were told that Cu has a 1+ charge, it will therefore have one more proton than it does electrons.

The mass number of an atom is defined as the number of protons plus the number of neutrons. Whenever you know the two of the three quantities in the equations shown below, you can determine the value for the third quantity.

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

For more details: See chapter 3 part 5 video or chapter 3 section 4 in the textbook.
3.11) Write the **formula** for each of the following compounds:

a) potassium bicarbonate

b) sodium bromide

c) iron(III) fluoride

d) sodium carbonate

e) iron(II) sulfate

f) barium hydroxide
3.11) Write the **formula** for each of the following compounds:

a) potassium bicarbonate

b) sodium bromide

c) iron(III) fluoride

d) sodium carbonate

e) iron(II) sulfate

f) barium hydroxide

**HINT:** Use the diagram below to find the formula of a compound when you are given its name.

**Binary Covalent (Molecular) Compound:**

1) Write the symbol of the first element in the compound’s name, then the symbol of the second element in the compound’s name.

2) Indicate how many atoms of each element the molecule contains using subscripts after the atomic symbol.

- The numbers of atoms are given in the molecule’s name in Greek prefixes

  - NOTE: if there is no Greek prefix in front of the first element in the name that implies the number is 1.

**Ionic Compound:**

1) Write the symbol/formula of the first ion in the compound’s name, then the symbol/formula of the second ion in the compound’s name.

2) Indicate the ratio of the ions in the compound using subscripts after each ion.

- The ratio of the ions is deduced by balancing the charges of the ions.

  - IMPORTANT: When there is more than one of a polyatomic ion in the formula unit we use parenthesis.

  *Example Mg(NO$_3$)$_2$*

For more help: See [chapter 3 part 10 video](#) or [chapter 3 section 7 in the textbook](#)
3.11) Write the formula for each of the following compounds:

a) potassium bicarbonate \( \text{KHCO}_3 \)
   (Note that bicarbonate is a polyatomic ion; see the table of polyatomic ions to get its formula and charge.)

b) sodium bromide \( \text{NaBr} \)

c) iron(III) fluoride \( \text{FeF}_3 \)
   • From the (III) in its name, we know the iron has a 3+ charge

d) sodium carbonate \( \text{Na}_2\text{CO}_3 \)

e) iron(II) sulfate \( \text{FeSO}_4 \)
   • We do not write \( \text{Fe}_2(\text{SO}_4)_2 \), we write the lowest ratio, 2:2 is equal to 1:1

f) barium hydroxide \( \text{Ba}(\text{OH})_2 \)
   • Because hydroxide (OH) is a polyatomic ion and its subscript is greater than 1, we write its formula in parenthesis and the subscript is written after/outside of the parenthesis.

For more details: See chapter 3 part 10 video or chapter 3 section 7 in the textbook.
3.12) Write the names for the following compounds. Be sure to use roman numerals with cations that can occur with multiple charges.

a) \( \text{Fe}_2(\text{CO}_3)_3 \)

b) \( \text{Cu}(\text{OH})_2 \)

c) \( (\text{NH}_4)_2\text{SO}_4 \)

d) \( \text{LiNO}_3 \)

e) \( \text{Mg(NO}_3)_2 \)

f) \( \text{AgCl} \)

g) \( \text{Al(OH)}_3 \)

h) \( \text{CaSO}_4 \)

i) \( \text{FeS} \)

j) \( \text{PbCl}_2 \)
3.12) Write the names for the following compounds. Be sure to use roman numerals with cations that can occur with multiple charges.

a) Fe$_2$(CO$_3$)$_3$

b) Cu(OH)$_2$

c) (NH$_4$)$_2$SO$_4$
  • Note: NH$_4^+$ is a polyatomic cation

d) LiNO$_3$

e) Mg(NO$_3$)$_2$

f) AgCl

g) Al(OH)$_3$

h) CaSO$_4$

i) FeS

j) PbCl$_2$

For more help: See chapter 3 part 10 video or chapter 3 section 7 in the textbook.
3.12) Write the names for the following compounds. Be sure to use roman numerals with cations that can occur with multiple charges.

a) \( \text{Fe}_2(\text{CO}_3)_3 \) iron(III) carbonate

Iron (Fe) is one of the metals that can occur with varying charge, you will need to deduce the charge on the iron cation based on the three carbonates and their charge.

b) \( \text{Cu(OH)}_2 \) copper(II) hydroxide

Cu is one of the metals that can occur with varying charge.

c) \( (\text{NH}_4)_2\text{SO}_4 \) ammonium sulfate

- Note: \( \text{NH}_4^+ \) is a polyatomic cation

d) \( \text{LiNO}_3 \) lithium nitrate

e) \( \text{Mg(NO}_3)_2 \) magnesium nitrate

f) \( \text{AgCl} \) silver chloride

Silver (Ag) is a transition metal, however it is not one that can occur with varying charge, therefore you do not need to indicate the charge in parenthesis after the cation name.

g) \( \text{Al(OH)}_3 \) aluminum hydroxide

h) \( \text{CaSO}_4 \) calcium sulfate

i) \( \text{FeS} \) iron(II) sulfide

Iron (Fe) is one of the metals that can occur with varying charge.

j) \( \text{PbCl}_2 \) lead(II) chloride

Lead (Pb) is one of the metals that can occur with varying charge.

For more details: See chapter 3 part 10 video or chapter 3 section 7 in the textbook.
3.13) Which of the following atom pairs would most likely be connected by a **covalent bond**?

a) Na and O  
b) Na and Cl  
c) O and Cl  
d) K and F
3.13) Which of the following atom pairs would most likely be connected by a **covalent bond**?

a) Na and O  

b) Na and Cl  

c) O and Cl  

d) K and F  

**HINT:**

**Covalent bonding** results from:
- Combining **nonmetal ions** with **nonmetal ions**.

**Ionic bonding** (ionic compounds) results from:
- Combining **metal ions** with **nonmetal ions**.  
  or  
- Combining **polyatomic ions** with **other ions**.
3.13) Which of the following atom pairs would most likely be connected by a **covalent bond**?

a) Na and O  

b) Na and Cl  

c) O and Cl  

d) K and F

**EXPLANATION:**

**Covalent bonding** results from:

- Combining **nonmetal ions** with **nonmetal ions**.

**Ionic bonding** (ionic compounds) results from:

- Combining **metal ions** with **nonmetal ions**.
  - or
- Combining **polyatomic ions** with **other ions**.
3.14) Which ion has the same number of *total electrons* as the noble gas argon (Ar)?

a) F^-

b) Br^-

c) Mg^{2+}

d) Ca^{2+}
3.14) Which ion has the same number of *total electrons* as the noble gas argon (Ar)?

a) F⁻

b) Br⁻

HINT:
In this problem, we are concerned with the number of *total electrons*, NOT the number of valence electrons.

Determine the number of electrons in Argon.

For more help: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.

c) Mg²⁺

d) Ca²⁺

Determine the number of electrons in the ions by starting with the number of electrons that would be contained in the neutral atoms and then adding or subtracting electrons based on the charge.
3.14) Which ion has the same number of total electrons as the noble gas argon (Ar)?

a) $F^-$ A fluorine atom contains 9 electrons. A fluoride ion has a charge of 1$^-$ because it contains one extra electron. This gives a total of 10 electrons.

b) $Br^-$ A bromine atom contains 35 electrons. A bromide ion has a charge of 1$^-$ because it contains one extra electron. This gives a total of 36 electrons.

c) $Mg^{2+}$ A magnesium atom contains 12 electrons. A magnesium ion has a charge of 2$^+$ because it contains two fewer electrons (than the atom). This gives a total of 10 electrons.

d) $Ca^{2+}$ A calcium atom contains 20 electrons. A calcium ion has a charge of 2$^+$ because it contains two fewer electrons (than the atom). This gives a total of 18 electrons.

EXPLANATION: In this problem, we are concerned with the number of total electrons, NOT the number of valence electrons.

Argon has a total of 18 electrons.

The number of electrons in the ions is determined by considering the number of electrons that are contained in the neutral atoms and then adding or subtracting electrons based on the charge.

For more help: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.
3.15) Write the names of the following compounds:

a) S₂I₄
b) P₅F₈
c) N₂O₄
d) NBr₃
e) N₂O₅
f) PCl₃
g) H₂S
h) N₂O
3.15) Write the names of the following compounds:

a) $\text{S}_2\text{I}_4$

b) $\text{P}_5\text{F}_8$

c) $\text{N}_2\text{O}_4$

d) $\text{NBr}_3$

e) $\text{N}_2\text{O}_5$

f) $\text{PCl}_3$

g) $\text{H}_2\text{S}$

h) $\text{N}_2\text{O}$

For more help: See chapter 3 part 8 video or chapter 3 section 6 in the textbook.
3.15) Write the names of the following compounds:

a) $\text{S}_2\text{I}_4$  \hspace{0.5cm} \text{disulfur tetriodide}

b) $\text{P}_5\text{F}_8$  \hspace{0.5cm} \text{pentaphosphorus octafluoride}

c) $\text{N}_2\text{O}_4$  \hspace{0.5cm} \text{dinitrogen tetroxide}

d) $\text{NBr}_3$  \hspace{0.5cm} \text{nitrogen tribromide}

e) $\text{N}_2\text{O}_5$  \hspace{0.5cm} \text{dinitrogen pentoxide}

f) $\text{PCl}_3$  \hspace{0.5cm} \text{phosphorus trichloride}

g) $\text{H}_2\text{S}$  \hspace{0.5cm} \text{dihydrogen monosulfide}

h) $\text{N}_2\text{O}$  \hspace{0.5cm} \text{dinitrogen monoxide}

EXPLANATION: Use the diagram below to find the name of a compound when you are given its formula.

For more details: See chapter 3 part 8 video or chapter 3 section 6 in the textbook.
3.16) Write the molecular formulas of the following compounds:

a) disulfur tetrafluoride
b) carbon trioxide
c) nitrogen pentoxide
d) nitrogen tribromide
e) dinitrogen heptachloride
f) carbon tetrachloride
g) hydrogen monochloride
h) trihydrogen monophosphide
i) dihydrogen monoxide
3.16) Write the molecular formulas of the following compounds:

a) disulfur tetrafluoride
b) carbon trioxide
c) nitrogen pentoxide
d) nitrogen tribromide
e) dinitrogen heptachloride
f) carbon tetrachloride
g) hydrogen monochloride
h) trihydrogen monophosphide
i) dihydrogen monoxide

**HINT:** Use the diagram below to find the formula of a compound when you are given its name.

Determine if the Compound is **Binary Covalent (Molecular)** or **Ionic**:

Does the compound contain only two types of **nonmetal** elements?

Yes  \( \rightarrow \) **Binary Covalent (Molecular) Compound:**

1) Write the symbol of the first element in the compound’s name, then the symbol of the second element in the compound’s name.
2) Indicate how many atoms of each element the molecule contains using subscripts after the atomic symbol.
   • The numbers of atoms are given in the molecule’s name in Greek prefixes
   • NOTE: if there is no Greek prefix in front of the first element in the name that implies the number is 1.

**Ionic Compound:**

1) Write the symbol/formula of the first ion in the compound’s name, then the symbol/formula of the second ion in the compound’s name.
2) Indicate the ratio of the ions in the compound using subscripts after each ion.
   • The ratio of the ions is deduced by balancing the charges of the ions.
   • IMPORTANT: When there is more than one of a polyatomic ion in the formula unit we use parenthesis.
   Example \( \text{Mg}(\text{NO}_3)_2 \)

**For more help:** See chapter 3 part 8 video or chapter 3 section 6 in the textbook.
3.16) Write the molecular formulas of the following compounds:

a) disulfur tetrafluoride \( \text{S}_2\text{F}_4 \)
b) carbon trioxide \( \text{CO}_3 \)
c) nitrogen pentoxide \( \text{NO}_5 \)
d) nitrogen tribromide \( \text{NBr}_3 \)
e) dinitrogen heptachloride \( \text{N}_2\text{Cl}_7 \)
f) carbon tetrachloride \( \text{CCl}_4 \)
g) hydrogen monochloride \( \text{HCl} \)
h) trihydrogen monophosphide \( \text{H}_3\text{P} \)
i) dihydrogen monoxide \( \text{H}_2\text{O} \)

**EXPLANATION:** Use the diagram below to find the formula of a compound when you are given its name.

- **Binary Covalent (Molecular) Compound:**
  1. Write the symbol of the first element in the compound’s name, then the symbol of the second element in the compound’s name.
  2. Indicate how many atoms of each element the molecule contains using subscripts after the atomic symbol.
     - The numbers of atoms are given in the molecule’s name in Greek prefixes
     - NOTE: if there is no Greek prefix in front of the first element in the name that implies the number is 1.

- **Ionic Compound:**
  1. Write the symbol/formula of the first ion in the compound’s name, then the symbol/formula of the second ion in the compound’s name.
  2. Indicate the ratio of the ions in the compound using subscripts after each ion.
     - The ratio of the ions is deduced by balancing the charges of the ions.
     - IMPORTANT: When there is more than one of a polyatomic ion in the formula unit we use parenthesis. Example \( \text{Mg(NO}_3)_2 \)

**For more details:** See chapter 3 part 8 video or chapter 3 section 6 in the textbook.
3.17) Calculate the molar mass of each compound given below.

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

a) H$_2$S

b) iron(II) nitrate
3.17) Calculate the molar mass of each compound given below.

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

a) H$_2$S

**HINT**: We cannot get the molar mass of a compound *directly* from the periodic table as we did for atomic molar masses.

To calculate the molar mass of a compound we **add up** the atomic molar masses of **all** atoms in the chemical formula.

b) iron(II) nitrate

**For more help**: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.
3.17) Calculate the molar mass of each compound given below.

a) \( \text{H}_2\text{S} \quad 34.09 \text{ g/mole} \)

b) iron(II) nitrate \( 179.87 \text{ g/mole} \)
3.17) Calculate the molar mass of each compound given below.

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

a) H₂S  **34.09 g/mole**

<table>
<thead>
<tr>
<th>Atom</th>
<th># of Atoms</th>
<th>Atomic Molar Mass</th>
<th>Total</th>
</tr>
</thead>
<tbody>
<tr>
<td>sulfur</td>
<td>1</td>
<td>32.07 g/mole</td>
<td>32.07 g/mole</td>
</tr>
<tr>
<td>hydrogen</td>
<td>2</td>
<td>1.01 g/mole</td>
<td>2.02 g/mole</td>
</tr>
</tbody>
</table>

Molar Mass of H₂S = 34.09 g/mole

b) iron(II) nitrate  **Fe(NO₃)₂  179.87 g/mole**

Fe  

(one mole of iron(II) ions)  

two moles of nitrate ions contain:

• two moles of nitrogen  
• six (2 x 3) moles of oxygen

one mole of Fe:  1 x 24.31 g/mole = 55.85 g/mole  
two moles of N:  2 x 14.01 g/mole = 28.02 g/mole  
six moles of O:  6 x 16.00 g/mole = 96.00 g/mole

The molar mass of Fe(NO₃)₂ is 179.87 g/mole

**For more details:** See chapter 3 part 11 video or chapter 3 section 9 in the textbook.
3.18) How many moles are contained in 251 grams of Fe$_2$O$_3$?
3.18) How many moles are contained in 251 grams of Fe₂O₃?

**HINT:**
First determine the **molar mass** of Fe₂O₃.
Next, convert from grams to moles using the molar mass as a **conversion factor**.

The chart shown below can be helpful when doing gram/mole/molecule conversions.

For more help: See [chapter 3 part 11 video](#) or chapter 3 section 9 in the textbook.
3.18) How many moles are contained in 251 grams of Fe$_2$O$_3$?

ANSWER: 1.57 moles

Your answer should have three significant figures.

CLICK HERE to see the complete solution for this problem
3.18) How many moles are contained in 251 grams of Fe$_2$O$_3$?

First determine the molar mass of Fe$_2$O$_3$.

<table>
<thead>
<tr>
<th>Ion</th>
<th># of ions in the Formula Unit</th>
<th>Molar Mass of ion</th>
<th>Total</th>
</tr>
</thead>
<tbody>
<tr>
<td>iron (III)</td>
<td>2</td>
<td>55.85 g/mole</td>
<td>= 111.70 g/mole</td>
</tr>
<tr>
<td>oxide</td>
<td>3</td>
<td>16.00 g/mole</td>
<td>= 48.00 g/mole</td>
</tr>
<tr>
<td><strong>Molar Mass (Formula Mass) of Fe$_2$O$_3</strong></td>
<td></td>
<td><strong>= 159.70 g/mole</strong></td>
<td></td>
</tr>
</tbody>
</table>

Next, convert from grams to moles using the molar mass as a conversion factor.

The relationship between moles and grams is the molar mass.

\[
\frac{251 \text{ grams Fe}_2\text{O}_3}{1 \text{ mole Fe}_2\text{O}_3} = \frac{159.70 \text{ grams Fe}_2\text{O}_3}{1.57 \text{ moles Fe}_2\text{O}_3}
\]

**ANSWER:** 1.57 moles

Your answer should have three significant figures.

For more details: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.
3.19) How many grams are contained in 0.482 moles of \( \text{Na}_2\text{CO}_3 \)?
3.19) How many grams are contained in 0.482 moles of Na$_2$CO$_3$?

**HINT:**

First determine the **molar mass** of Na$_2$CO$_3$.

Next, convert from moles to grams using the molar mass as a **conversion factor**.

The chart shown below can be helpful when doing gram/mole/molecule conversions.

For more help: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.
3.19) How many grams are contained in 0.482 moles of Na$_2$CO$_3$?

ANSWER: 51.1 grams

Your answer should have three significant figures.
3.19) How many grams are contained in 0.482 moles of Na\(_2\)CO\(_3\)?

First determine the **molar mass** of Na\(_2\)CO\(_3\).

\[
\begin{align*}
\text{Na} & \quad \text{one mole of carbonate ions contains:} \\
\text{two moles of sodium ions} & \quad \cdot \text{one mole of carbon} \\
& \quad \cdot \text{three moles of oxygen} \\
\text{two moles of Na:} & \quad 2 \times 22.99 \text{ g/mole} = \text{45.98 g/mole} \\
\text{one mole of C:} & \quad 1 \times 12.01 \text{ g/mole} = \text{12.01 g/mole} \\
\text{three moles of O:} & \quad 3 \times 16.00 \text{ g/mole} = \text{48.00 g/mole} \\
\text{The molar mass of Na}_2\text{CO}_3 \text{ is } & \quad 105.99 \text{ g/mole}
\end{align*}
\]

Next, convert from moles to grams using the molar mass as a **conversion factor**.

\[
\text{0.482 moles Na}_2\text{CO}_3 \times \frac{105.99 \text{ grams Na}_2\text{CO}_3}{1 \text{ mole Na}_2\text{CO}_3} = 51.1 \text{ grams Na}_2\text{CO}_3
\]

**ANSWER:** 51.1 grams

Your answer should have **three** significant figures.

For more details: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.
3.20) What is the mass (in grams) of $9.39 \times 10^{24}$ molecules of methanol (CH$_4$O)?
3.20) What is the mass (in grams) of $9.39 \times 10^{24}$ \textit{molecules} of methanol (CH$_4$O)?

\textbf{HINT:}

First determine the \textbf{molar mass} of CH$_4$O.

Next, convert from \textbf{molecules} to \textbf{grams}. This is a \textbf{TWO STEP unit conversion problem}.

The chart shown below can be helpful when doing gram/mole/molecule conversions.

\textbf{For more help:} See chapter 3 part 11 video or chapter 3 section 9 in the textbook.
3.20) What is the mass (in grams) of $9.39 \times 10^{24}$ molecules of methanol (CH$_4$O)?

**ANSWER:** 500. grams CH$_4$O (or $5.00 \times 10^2$ grams CH$_4$O)

Your answer should have three significant figures.
3.20) What is the mass (in grams) of \(9.39 \times 10^{24}\) **molecules** of methanol \((\text{CH}_4\text{O})\)?

**ANSWER:** 500. grams \(\text{CH}_4\text{O}\)
(or \(5.00 \times 10^2\) grams \(\text{CH}_4\text{O}\))

Your answer should have **three** significant figures.

Next, convert from **molecules** to **grams**. This is a **TWO STEP** unit conversion problem.

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

\[
\begin{array}{c|c|c}
9.39 \times 10^{24} \text{ CH}_4\text{O} \text{ molecules} & 1 \text{ mole CH}_4\text{O} & = 15.5928263 \text{ CH}_4\text{O} \text{ molecules} \\
6.022 \times 10^{23} \text{ CH}_4\text{O} \text{ molecules} & \text{ (unrounded)}
\end{array}
\]

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

\[
\begin{array}{c|c|c}
15.5928263 \text{ CH}_4\text{O} \text{ molecules} & 32.05 \text{ grams CH}_4\text{O} & = 500. \text{ grams CH}_4\text{O} \\
1 \text{ mole CH}_4\text{O} & \text{ (or } 5.00 \times 10^2 \text{ grams CH}_4\text{O})
\end{array}
\]

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

\[
\begin{array}{c|c|c|c}
9.39 \times 10^{24} \text{ CH}_4\text{O} \text{ molecules} & 1 \text{ mole CH}_4\text{O} & 32.05 \text{ grams CH}_4\text{O} & = 500. \text{ grams CH}_4\text{O} \\
6.022 \times 10^{23} \text{ CH}_4\text{O} \text{ molecules} & \text{ (or } 5.00 \times 10^2 \text{ grams CH}_4\text{O}) & 1 \text{ mole CH}_4\text{O}
\end{array}
\]

For more details: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.
What is the mass (in grams) of $9.43 \times 10^{23}$ molecules of sucrose ($C_{12}H_{22}O_{11}$)?
3.21) What is the mass (in grams) of $9.43 \times 10^{23}$ molecules of sucrose ($C_{12}H_{22}O_{11}$)?

**HINT:**

First determine the **molar mass** of $C_{12}H_{22}O_{11}$.

Next, convert from **molecules** to **grams**. This is a **TWO STEP unit conversion problem**.

The chart shown below can be helpful when doing gram/mole/molecule conversions.

**For more help:** See chapter 3 part 11 video or chapter 3 section 9 in the textbook.
3.21) What is the mass (in grams) of $9.43 \times 10^{23}$ molecules of sucrose ($C_{12}H_{22}O_{11}$)?

**ANSWER:** 536 grams $C_{12}H_{22}O_{11}$

Your answer should have **three** significant figures.
3.21) What is the mass (in grams) of 9.43 x 10^{23} molecules of sucrose (C_{12}H_{22}O_{11})?

**First determine the molar mass of C_{12}H_{22}O_{11}.

{12 \text{ moles of C}: 12 \times 12.01 \text{ g/mole} = 144.12 \text{ g/mole}}

{22 \text{ moles of H}: 22 \times 1.01 \text{ g/mole} = 22.22 \text{ g/mole}}

{11 \text{ moles of O}: 11 \times 16.00 \text{ g/mole} = +176.00 \text{ g/mole}}

The molar mass of C_{12}H_{22}O_{11} is 342.34 \text{ g/mole}

Next, convert from molecules to grams. This is a TWO STEP unit conversion problem.

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

\[
\begin{array}{c|c|c}
9.43 \times 10^{23} \text{ C}_{12}\text{H}_{22}\text{O}_{11} \text{ molecules} & 1 \text{ mole C}_{12}\text{H}_{22}\text{O}_{11} & = 1.565924942 \text{ C}_{12}\text{H}_{22}\text{O}_{11} \text{ molecules (unrounded)} \\
6.022 \times 10^{23} \text{ C}_{12}\text{H}_{22}\text{O}_{11} \text{ molecules} & \end{array}
\]

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

\[
\begin{array}{c|c|c}
1.565924942 \text{ C}_{12}\text{H}_{22}\text{O}_{11} \text{ molecules} & 342.34 \text{ grams C}_{12}\text{H}_{22}\text{O}_{11} & = 536 \text{ grams C}_{12}\text{H}_{22}\text{O}_{11} \\
1 \text{ mole C}_{12}\text{H}_{22}\text{O}_{11} & \end{array}
\]

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

\[
\begin{array}{c|c|c}
9.43 \times 10^{23} \text{ C}_{12}\text{H}_{22}\text{O}_{11} \text{ molecules} & 1 \text{ mole C}_{12}\text{H}_{22}\text{O}_{11} & 342.34 \text{ grams C}_{12}\text{H}_{22}\text{O}_{11} \\
6.022 \times 10^{23} \text{ C}_{12}\text{H}_{22}\text{O}_{11} \text{ molecules} & 1 \text{ mole C}_{12}\text{H}_{22}\text{O}_{11} & \end{array}
\]

ANSWER: 536 grams C_{12}H_{22}O_{11}

Your answer should have **three** significant figures.

For more details: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.
3.22) How many molecules are contained in 38.33 grams of NO₂?
3.22) How many molecules are contained in 38.33 grams of NO₂?

**HINT:**

First determine the **molar mass** of NO₂.

Next, convert from **grams** to **molecules**. This is a **TWO STEP unit conversion problem**.

The chart shown below can be helpful when doing gram/mole/molecule conversions.

For more help: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.
3.22) How many molecules are contained in 38.33 grams of NO₂?

Answer: \(5.017 \times 10^{23}\) NO₂ molecules

Your answer should have four significant figures.
3.22) How many molecules are contained in 38.33 grams of NO₂?

First determine the molar mass of NO₂.

\[
\text{one mole of N: } 1 \times 14.01 \text{ g/mole} = 14.01 \text{ g/mole}
\]
\[
\text{two moles of O: } 2 \times 16.00 \text{ g/mole} = 32.00 \text{ g/mole}
\]

The molar mass of NO₂ is 46.01 g/mole

Next, convert from grams to molecules. This is a TWO STEP unit conversion problem.

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

\[
\frac{38.33 \text{ grams NO}_2}{46.01 \text{ grams NO}_2} = 0.8330797653 \text{ moles NO}_2 \quad \text{(unrounded)}
\]

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

\[
\frac{0.8330797653 \text{ moles NO}_2}{1 \text{ mole NO}_2} = 6.022 \times 10^{23} \text{ NO}_2 \text{ molecules} = 5.017 \times 10^{23} \text{ NO}_2 \text{ molecules}
\]

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

\[
\frac{38.33 \text{ grams NO}_2}{46.01 \text{ grams NO}_2} = \frac{6.022 \times 10^{23} \text{ NO}_2 \text{ molecules}}{1 \text{ mole NO}_2} = 5.017 \times 10^{23} \text{ NO}_2 \text{ molecules}
\]

**ANSWER:** 5.017 \times 10^{23} \text{ NO}_2 \text{ molecules}

Your answer should have **four** significant figures.

For more details: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.
3.23) How many molecules are contained in 25.23 grams of dihydrogen sulfide (H₂S)?
3.23) How many molecules are contained in 25.23 grams of dihydrogen sulfide (H₂S)?

**HINT:**
First determine the **molar mass** of NO₂.
Next, convert from **grams** to **molecules**. This is a **TWO STEP unit conversion problem**.
The chart shown below can be helpful when doing gram/mole/molecule conversions.

For more help: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.
3.23) How many molecules are contained in 25.23 grams of dihydrogen sulfide (H₂S)?

**ANSWER:** \(4.457 \times 10^{23}\) H₂S molecules

Your answer should have **four** significant figures.

This is the last chapter 3 review problem.
3.23) How many molecules are contained in 25.23 grams of dihydrogen sulfide (H$_2$S)?

ANSWER: $4.457 \times 10^{23}$ H$_2$S molecules.

First determine the **molar mass** of H$_2$S.

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

\[
\begin{align*}
\text{two moles of } \text{H:} & \quad 2 \times 1.01 \text{ g/mole} = 2.02 \text{ g/mole} \\
\text{one mole of } \text{S:} & \quad 1 \times 32.07 \text{ g/mole} = 32.07 \text{ g/mole}
\end{align*}
\]

The molar mass of H$_2$S is **34.09 g/mole**.

Next, convert from **grams** to **molecules**. This is a **TWO STEP unit conversion problem**.

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

\[
\begin{align*}
& \quad 25.23 \text{ grams H}_2\text{S} \\
\text{1 mole H}_2\text{S} & \quad 34.09 \text{ grams H}_2\text{S}
\end{align*}
\]

\[
= 0.8330797653 \text{ moles H}_2\text{S}
\]

(unrounded)

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

\[
\begin{align*}
& \quad 0.8330797653 \text{ moles H}_2\text{S} \\
6.022 \times 10^{23} \text{ H}_2\text{S molecules} & \quad 1 \text{ mole H}_2\text{S}
\end{align*}
\]

\[
= 4.457 \times 10^{23} \text{ H}_2\text{S molecules}
\]

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

\[
\begin{align*}
& \quad 25.23 \text{ grams H}_2\text{S} \\
\text{1 mole H}_2\text{S} & \quad 34.09 \text{ grams H}_2\text{S}
\end{align*}
\]

\[
\begin{align*}
& \quad 1 \text{ mole H}_2\text{S} \\
6.022 \times 10^{23} \text{ H}_2\text{S molecules} & \quad 1 \text{ mole H}_2\text{S}
\end{align*}
\]

\[
= 4.457 \times 10^{23} \text{ H}_2\text{S molecules}
\]

This is the last chapter 3 review problem.