Chemistry 2 Final Study Guide

Compiled by Adam Furman

Jump To

Test Overview

The test consists of 70 multiple-choice questions presented on a Scantron sheet. Each question is worth 2 points for a total of 140 points overall. The final test is around 12% of the semester final grade in the class.

It covers topics from textbook chapters 5, 6, 19.1, 7, 8, and 9.
The following items will be made available for use on the exam by Mr. Crider:

- Periodic Table
- Activity Series
- Solubility Rules (no polyatomic ions listed)
- Electronegativity Chart

In addition, you may use the following of your own resources on the test:

- Calculator
- 4x6in Notecard (half of a regular sheet of paper)
- #2 or HB Mechanical Pencil

You must also bring your Chemistry textbook to turn in during the test.

Note: You should also prepare yourself mentally and physically for a strenuous final. [1]

(Revision 2) Changes Highlighted *Thanks to Nitin and Abby*
Newest revisions: https://furmada.github.io/Homepage/chem_studyguide_latest.pdf

This guide is intended to provide an overview of both final exam testing material as well as review topics. Unless otherwise noted, all answers are computed to significant figures.

If you find an error in this guide, please email me at furmada@gmail.com.

[1] A good list of pre-test practices is available here.

Conversions

Conversions allow you to change between forms of measurement. At their core, they are a simple algebraic concept which can be extended to higher complexity.

Given a value in one system of measurement, one can convert to another form of measurement by using a *conversion ratio*. A conversion ratio can be set up as a fraction, where the numerator is the desired new unit, and the denominator is the current unit. The number 1 is always placed next to the unit that is larger. For example, the 1 is always placed next to the unit mole. The other unit then receives the conversion *constant*, that is, a number representing the relationship between the two units. For example, Avogadro's Number is a constant.

$$Original\ Value\ (Original\ Unit) \times \frac{Desired\ Unit}{Original\ Unit} = Resultant\ Value\ (Desired\ Unit)$$

This ratio can be simplified into the mnemonic statement “want over have.”

Example: Meters to Centimeters

→*Convert 3.00m to centimeters.*

Set up the conversion ratio, starting with the original value.

$$3.00m \times \frac{100.\,cm}{1m} =$$

Notice that the *m* for meters and the *m* in the denominator cancel out, resulting in the desired unit of centimeters. Then, simply multiply the two numbers together.

$$3.00m \times \frac{100.\,cm}{1m} = 300.\,cm$$

Example: Moles to Atoms

→*Convert 2.0 moles of a substance to atoms.*

As before, set up the conversion ratio first. Then multiply the numbers together, ensuring that the units have cancelled out. The constant at use here is Avogadro’s Number, which equals 6.022×10^{23}.

$$2.0\cancel{mol} \times \frac{6.022 \times 10^{23} atoms}{1\cancel{mol}} = 1.2 \times 10^{24} atoms$$

Significant Figures

Significant Figures are a set of rules that is used to determine the precision to which an answer should be given. Simply put, the final answer to a problem should only be as precise as its least reliable measurement.

Determining how many significant figures a number has uses the following rules:

- All non-zero digits are significant.
- All zeros in between significant digits are also significant.
- All zeros after both the decimal dot and a non-zero digit are significant.
- Leading zeros are never significant.

When adding or subtracting two numbers, the answer must be rounded to the least amount of decimal places. When multiplying or dividing, the answer is rounded to the least amount of significant figures.

Rounding only uses the next digit, numbers are not rounded starting at the right end.

Example: Determining Significance
→*Determine the number of significant figures in the number* 0.0120

Remember that leading zeros are never significant, but trailing zeros are, as long as they are behind both a non-zero digit and the decimal place. This number then has 3 significant figures.

Example: Significant Figure Addition
→*Add the numbers* 10. *and* 0.5

First, add the numbers to get 10.5. Then, determine the number of decimal places in each. 10. has none, but 0.5 has one. Therefore, the answer must have no decimals. But 0.5 rounds up, so the answer is 11.

Example: Significant Figure Multiplication
→*Multiply* 2.5 *by* 10

First, multiply to get 25. Then, determine the significance of each number. 2.5 has two significant figures, but 10 only has one! This means that the answer must also only have one significant figure. The correct answer is therefore 30

Polyatomic Ions

Polyatomic ions are ions consisting of more than one atom joined together in a covalent bond. This bond requires the addition or removal of one or more electrons to form, which means that the whole covalent structure acts like a single-atom ion in a chemical reaction and forms bonds based on its oxidation number.

You must memorize the common polyatomic ions and their charges for use in chemical reactions.

The name of a polyatomic ion depends on how many oxygen atoms it contains. The –ate ion is the base, the –ite ions have 1 less oxygen, the hypo- -ite ions have 2 less oxygen, and the per- -ate ions have 1 more oxygen.

Ions with a "Hydrogen" in their name have a charge that is 1 more than the –ate.

It is helpful to remember the following basic –ate ions:

- Acetate ($C_2H_3O_2^{-1}$)
- Carbonate (CO_3^{-2})
- Chlorate (ClO_3^{-1})
- Nitrate (NO_3^{-1})
- Phosphate (PO_4^{-3})
- Sulfate (SO_4^{-2})

Common Polyatomic Ions

Name	***Symbol***
Acetate	$C_2H_3O_2^{-1}$
Ammonium	NH_4^{+1}
Carbonate	CO_3^{-2}
Chlorate	ClO_3^{-1}
Chlorite	ClO_2^{-1}
Chromate	CrO_4^{-2}
Cyanide	CN^{-1}
DH Phosphate	$H_2PO_4^{-1}$
Dichromate	$Cr_2O_7^{-2}$
H. Carbonate (Bicarbonate)	HCO_3^{-1}
H. Phosphate	HPO_4^{-1}
H. Sulfate	HSO_4^{-1}
H. Sulfite	HSO_3^{-1}
Hydroxide	OH^{-1}
Hypochlorite	ClO_1^{-1}
Nitrate	NO_3^{-1}
Nitrite	NO_2^{-1}
Perchlorate	ClO_4^{-1}
Permanganate	MnO_4^{-1}
Peroxide	O_2^{-2}
Phosphate	PO_4^{-3}
Phosphite	PO_3^{-3}
Sulfate	SO_4^{-2}
Sulfite	SO_3^{-2}

Example: Polyatomic Bond

→*What ionic compound is formed by combining Sodium and Phosphate?*

First, write the reactants and determine their oxidation numbers.

$$Na^{+1}, {PO_4}^{-3}$$

Then, combine the two, balancing the charges by adding a subscript of 3 to the Sodium ion.

$$Na_3(PO_4)$$

Acid Naming

An acid is a compound in which the cation (positively charged atom) is replaced by one or more Hydrogen atoms. They can form from single elements or polyatomic ions.

Acid names follow these conventions:

- Acids formed from a polyatomic –ate have names ending in –ic.
- Acids formed from a polyatomic –ite have names ending in –ous.
- Acids formed from elemental ions have the prefix **hydro**- and the ending –ic.

Example: Name an Acid

→*What is the name of the acid* H_3PO_4*?*

Notice that the base compound for this acid is Phosphate, PO_4^{-3}. This is an –ate polyatomic ion, so then the acid is called "phosphic acid." Note, this also has an alternate name: "phosphoric acid." Both are acceptable on the test.

Example: Name an Acid (Again)

→*What is the name of the compound* HCl*?*

Notice that the base compound for the acid is Chlorine, Cl^{-1}. This is a standalone element, so the name of the acid is "hydrochloric acid."

Example: Acid Formula

→*Write the chemical formula for nitrous acid.*

This acid ends in –ous, which means it was formed from a polyatomic –ite. The only polyatomic ion with the "nitr-" beginning that is an –ite is Nitrite, NO_2^{-1}. Therefore, the formula of this acid is HNO_2.

Example: Acid Formula (Again)

→*Write the chemical formula for carbonic acid.*

This acid ends in –ic, which means it was formed from a polyatomic –ate. The only polyatomic ion with the "carb-" beginning that is an –ate is Carbonate, CO_3^{-2}. Therefore, the formula of this acid is H_2CO_3.

Bond Polarity

The polarity of a bond is determined by its electronegativity. Electronegativity is the ability of a molecule to attract electrons to itself. The highest electronegativity can be found at the upper right of the periodic table (Fluorine), and it decreases going down and to the left, with Francium having the lowest electronegativity.

The type of bond can be found by calculating the dipole moment between its components. The dipole moment is the absolute difference between the electronegativity values of the two elements in the bond.

$$DM = |E_0 - E_1|$$

The following are the types of bonds and what dipole moment is associated with each:

- Nonpolar Covalent ($DM \leq 0.5$)
  - This bond's atoms share electrons equally. An example of this is diatomic chlorine, Cl_2.
- Polar Covalent ($0.5 < DM \leq 2.1$)
  - The electrons in this bond are shared unequally, with the more electronegative atom snagging more of them. An example is carbon monochlorine, CCl.
- Ionic ($DM > 2.1$)
  - All of the electrons are taken by one of the atoms. Sodium Chloride, NaCl, is an example.

Example: Dipole Moment

→*What is the dipole moment in the bond between H and O in H_2O?*

First, find the electronegativity of H and O. It is 2.2 for H and 3.44 for O. Now, use the formula to subtract and find the dipole moment.

$$|2.2 - 3.44| = 1.2$$

Example: Bond Type

→*What type of bond is the one between H and O in H_2O?*

Using the electronegativity of 1.2 from the previous example and the classification chart above, this bond is polar covalent.

Ionic Bonds

An ionic bond is formed when two oppositely-charged molecules, called ions, exchange electrons. The atoms on the left side of the periodic table (groups 1 and 2) are called *cations*, and they have positive oxidation numbers. They accept electrons in an ionic bond because they contain more protons than electrons. The atoms on the right side of the table (groups 13 through 17) are *anions*, and they have negative oxidation numbers. They will "give away" electrons in an ionic bond because they have more of them than of protons. Ionic bonds only occur between a metal and a nonmetal.

In order for an ionic bond to work, the positive and negative charges must balance out exactly, so that the overall charge on the new molecule is 0.

Transition metals will change their oxidation number to suit different reactions. Their charge is represented using roman numerals next to their name. For example, Iron(II) is a Fe atom with a charge of +2.

There are several exceptions to the transition metal rule:

- Cadmium (Cd) has a charge of +2
- Zinc (Zn) has a charge of +2
- Silver (Ag) has a charge of +1

On the other hand, Tin (Sn), Lead (Pb), and Bismuth (Bi) are non-transition metals that change their oxidation numbers like transition metals.

When naming ionic compounds, add the –ide suffix to the last element if it is an element and not a polyatomic. For example, Lithium Sulf**ide** is the name for Li_2S. If the anion is a polyatomic, simply write its name without adding the suffix.

Example: Naming an Ionic Compound

→*What is the name for the ionic compound* Mg_3N_2*?*

This compound is named Magnesium Nitride.

Example: Naming an Ionic Compound (Again)

→*What is the name for the ionic compound* K_2SO_4*?*

This compound is named Potassium Sulfate.

Example: Naming an Ionic Compound (Once More)

→*What is the name for the ionic compound* FeO*?*

Note that you are given a transition metal. In this case, you must determine its charge. To do so, simply find the other ion's oxidation number.

This compound is named Iron(II) Oxide.

The Octet Rule

The octet rule describes how atoms fill their valence shells with electrons. Most elements wish to have 8 electrons in their outer **s** and **p** orbitals. This means that they wish to adopt the valence shell of a noble gas.

The octet rule determines an element's oxidation number. A negative number means an element is missing that many electrons to complete its octet. A positive number means the element will give up that number of electrons to achieve the same goal.

The octet rule has the following exceptions:

- Hydrogen (H) completes a duet, it requires only 2 electrons.
- Beryllium (Be) completes a quartet, it requires 4 electrons.
- Boron (B) and Aluminum (Al) complete sextets, they require 6 electrons.

Example: Complete the Octet

→*How many electrons does Phosphorous need to fill its valence shell?*

Phosphorous has an oxidation number of -3. It requires 3 electrons to complete the octet.

Example: Complete the Octet (Again)

→*How many electrons does Beryllium need to fill its valence shell?*

Beryllium has an oxidation number of +2. However, it is an exception to the octet rule. It requires 2 more electrons to fill its valence shell.

Time for a Break

Congratulations! You are well on your way to being prepared for the Chemistry 2 final exam. Have some chemistry jokes as a reward.

- I have so many chemistry jokes. I'm just afraid they won't get a good reaction!
- Why did the noble gas cry? Because all his friends Argon.
- How can you spot a chemist in the restroom? They wash their hands before they go.
- What is a cation afraid of? A dogion.
- A sign on the chemistry hotel reads: Great day rates, even better NO_3^{-1}s!
- The name's Bond. Ionic Bond. Taken, not shared.
- Are you made of Beryllium, Gold, and Titanium? Because you're BeAuTi-full!
- I would put more chemistry jokes here, but all the good ones Argon.

Covalent Bonds

A covalent bond occurs when two nonmetals share electrons. As discussed previously in the section on bond polarity, a covalent bond can either be nonpolar and share the electrons evenly, or polar and have the electrons shift to one atom. Covalent bonds complete the octet for the atoms involved by sharing the electrons in between the two molecules. A covalent bond can be single, double, or triple depending on how many pairs of electrons are shared (up to three pairs of 2 electrons, or 6 electrons total).

Covalent compounds are named depending on how many of each atom they contain. The following are the prefixes used:

- mono- (one)
- di- (two)
- tri- (three)
- tetr(a)- (four)
- pent(a)- (five)
- sext(a)- (six)
- sept(a)- (seven)
- oct(a)- (eight)
- non(a)- (nine)
- dec(a)- (ten)

The only exception is that if there is only one of the first element, the mono- prefix is not written.

Diatomic Elements

These elements occur in nature in covalently-bonded pairs.

- Nitrogen (N)
- Oxygen (O)
- Fluorine (F)
- Chlorine (Cl)
- Bromine (Br)
- Iodine (I)
- Hydrogen (H)

Example: Naming a Covalent Compound

→*What is the name for the covalent compound* CF_4*?*

The name of this compound is carbon tetrafluoride.

Example: Naming a Covalent Compound (Again)

→*What is the name for the covalent compound* N_2O_5*?*

The name of this compound is dinitrogen pentoxide.

Example: Covalent Compound Formula

→*Write the formula for boron septiodide*.

The formula is BI_7.

Lewis Structures

A Lewis Structure is a visual depiction of a molecule's bonds and electrons. In the diagram, the least electronegative atom is in the center, with the others spaced out evenly around it. To draw a diagram, follow these steps[2]:

- Count all the valence electrons, multiplying by any subscripts. Also include the charge of the molecule (if it is given).
- Draw the atoms, with the least electronegative in the center (or, if present, Carbon (C)).
    - Note that Hydrogen (H) is never the central atom.
- Draw single bonds to the central atom.
- Complete the octet (or equivalent) for all the atoms, starting with the central one.
    - If an octet cannot be completed, change one or more of the single bonds to a double bond.
- If the molecule had a charge, enclose the diagram in [brackets] and write the charge on the upper-right.

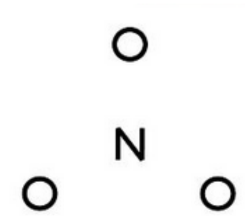
In some cases, a double bond is drawn to one of many correct possibilities. This is called resonance, and to correctly represent it, use a dashed line by any possible location of the bond.

Example: Draw the Lewis Structure

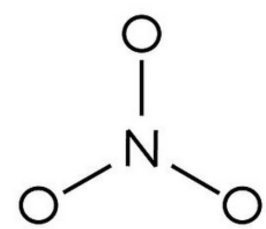
→*Draw the Lewis Structure for the molecule* NO_3^{-1}, Nitrate.

Start by counting the valence electrons. Nitrogen has 5, the three oxygens together have 18, and the overall charge is negative, which means that electrons are added (1 in this case). The total number of all valence electrons is 24.

Now, draw the basic structure.

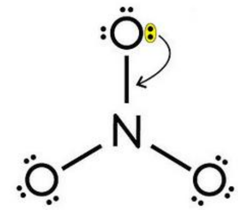


Draw single bonds.

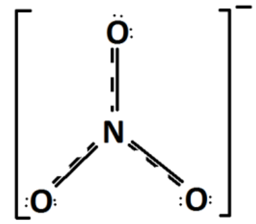


[2] The steps are taken from this infographic.

Complete the octets. Note that Nitrogen is incomplete, so a double bond must be present.



The structure is resonant, because each one of the oxygen atoms can be the location of the double bond.



Because of the difficulty of making digital Lewis structures, further practice is omitted here.

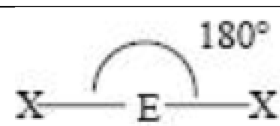
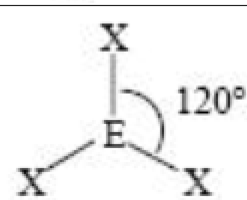
However, there are many resources on the internet, such as Wolfram | Alpha which can help you study Lewis structures.

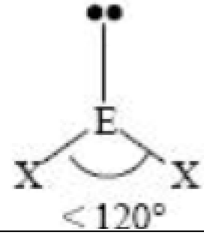
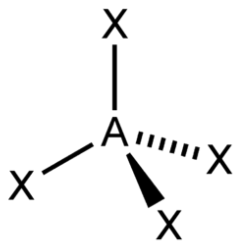
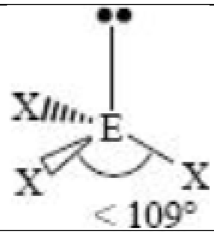
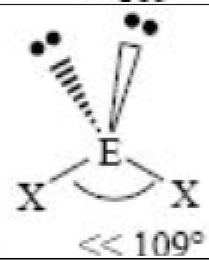
Three-Dimensional Diagrams

Molecules can be represented in three dimensions using VSEPR (Valence Shell Electron Pair Repulsion) diagrams. These diagram represent how a molecule would look in "real life." They are categorized by shape, hybridization (one less than the number of bonds the central atom makes, a value also known as the steric number), bond angle, and polarity.

When drawing the shape, the filled-out black triangle bond indicates an atom that is coming "out of" the paper and towards the observer. The dashed triangle is going "into" the paper.

The following is a table of VSEPR shapes and their data.

Shape Name	**Shape**	**Bond Ang.**	**Hybr.**	**Polarity**
Linear		180	sp	NP
Trigonal Planar		120	sp2	NP

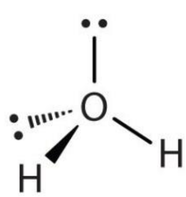
Bent (Steric #3)		120	sp2	P
Tetrahedral		109	sp3	NP
Pyramidal		109	sp3	P
Bent (Steric #4)		109	sp3	P

Note that VSEPR shapes do not represent resonance. Draw the double bond of a resonant structure on the atom of your choosing.

Example: Draw the VSEPR Diagram

→*Draw the VSEPR Diagram for water (*H_2O*).*

Following the steps for drawing a Lewis structure, count the valence electrons and draw the basic shape. Complete the octet for Oxygen by adding two pairs of electrons. As a result, the diagram will have a Bent (Steric #4) shape, with a bond angle of 109 and sp3 hybridization.



Hydrocarbons

A hydrocarbon is a large, complex molecule consisting of carbon and hydrogen. These molecules are divided in naming by what the highest bond is (single, double, or triple). To identify the hydrocarbon, first find the longest continual chain of carbons. From that point, the branches are called *functional groups*, and they are numbered from the closest end of the carbon chain.

The three categories, along with their formulas, are:

- Alkane hydrocarbons, with the formula $H = 2C + 2$ where C is the number of carbon atoms.
- Alkene hydrocarbons, with the formula $H = 2C$.
- Alkyne hydrocarbons, with the formula $H = 2C - 2$.

The name of a hydrocarbon depends on how many carbon atoms are in its longest chain. If the hydrocarbon is an –ene or an –yne, the location of the double bond is indicated in the name.

The prefixes are below:

- 1C: meth-
- 2C: eth-
- 3C: prop-
- 4C: but-
- 5C: pent-
- 6C: hex-
- 7C: hept-
- 8C: oct-
- 9C: non-
- 10C: dec-

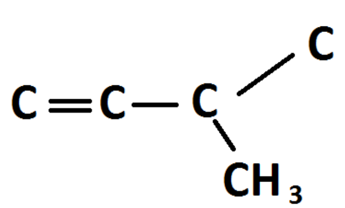
For example, a molecule with 3 carbons and 8 hydrogens would be called "propane."

A molecule with 6 carbons and 10 hydrogens would be "hexyne."

If one or more of the hydrogen atoms is replaced with a hydroxide (OH) molecule, the suffix –ol is added, turning the hydrocarbon into an alcohol.

Example: Name the Hydrocarbon

→*What is the name of the hydrocarbon pictured below?*



This hydrocarbon is 3 methyl 1 butene.

The longest chain has 4 carbons. It has a double bond at index 1 and a functional group at index 3.

Chemical Reactions

A chemical reaction is the process by which various substances interact to form new compounds. In it, the *reactants*, written on the left side of the equation, yield *products* on the right side. A reaction is considered to be balanced when there is an equal amount of each element on both sides of the equation.

There are six types of chemical reactions:

Name	**General Equation**
Combustion	$C_nH_m + O_2 \xrightarrow{yields} H_2O + CO_2$
Synthesis	$A + B \xrightarrow{yields} AB$
Decomposition	$AB \xrightarrow{yields} A + B$
Single Displacement	$AB + C \xrightarrow{yields} CB + A$
Double Displacement	$AB + CD \xrightarrow{yields} AD + CB$
Acid/Base Double Disp.	$HA + B(OH) \xrightarrow{yields} H_2O + BA$

Single displacement reactions will only react if the reactant (denoted as *C*) above, is higher on the activity series than *A* (or, if *C* is a halogen, *B*). You will be provided an activity series on the test, so it is not listed here.

For double displacement reactions, you must determine the solubility (phase) of each of the reactants and products.

- *(aq)* stands for aqueous, or dissolved in water.
- *(s)* stands for solid.
- *(l)* stands for liquid.
- *(g)* stands for gaseous.

If all the products are aqueous, then no reaction has occurred. Balance the reaction only after determining solubility. Solubility rules will be provided on the test, so they are not listed here.

Example: Identify the Reaction Type

→*Identify the following the reactions as one of the six types:* $CaO + Li \xrightarrow{yields} Li_2O + Ca.$

Notice that one element, Calcium, gets "pushed out" of its bond with oxygen by the more reactive Lithium. This is a single displacement reaction.

Net Ionic Equations

A net ionic equation represents the ionic bond that occurs in a double displacement reaction. It shows which ions interact to form the precipitate.

To find the net ionic equation, you must first find the total ionic equation. Simply break apart all the molecules, changing subscripts to coefficients and adding charges. Do not break apart the precipitate, keep it together and do not write a charge.

Once you have found the total ionic equation, identify the ions which are not a part of the precipitate and eliminate them from both sides.

Example: Net Ionic Equation

→*Find the net ionic equation for the following reaction:*

$$Na_2SO_{4\,(aq)} + Ca(NO_3)_{2\,(aq)} \xrightarrow{yields} CaSO_{4\,(s)} + 2NaNO_{3\,(aq)}$$

First, write the total ionic equation.

$$2Na^+{}_{(aq)} + SO_4{}^{-2}{}_{(aq)} + Ca^{+2}{}_{(aq)} + 2NO_3{}^-{}_{(aq)} \xrightarrow{yields} CaSO_{4\,(s)} + 2Na^+{}_{(aq)} + 2NO_3{}^-{}_{(aq)}$$

Now, identify and eliminate the ions that are the same on both sides.

$$\cancel{2Na^+{}_{(aq)}} + SO_4{}^{-2}{}_{(aq)} + Ca^{+2}{}_{(aq)} + \cancel{2NO_3{}^-{}_{(aq)}} \xrightarrow{yields} CaSO_{4\,(s)} + \cancel{2Na^+{}_{(aq)}} + \cancel{2NO_3{}^-{}_{(aq)}}$$

Finally, rewrite to get the net ionic equation.

$$SO_4{}^{-2}{}_{(aq)} + Ca^{+2}{}_{(aq)} \xrightarrow{yields} CaSO_{4\,(s)}$$

Empirical and Molecular Formulas

The empirical formula is a representation of the formula of a substance given the percent composition of that substance. To find the empirical formula, assume that you are given a 100.g sample, and then find how many moles of each component you have. Then, divide each by the lowest amount of moles to find the subscript for each. If the division operation results in a number ending in .5, double all of the values to arrive at whole numbers.

The molecular formula is the actual representation of the formula for a given sample size. Once you have found the empirical formula, find the ratio of the given molecular molar mass to the mass of the empirical formula. Then, multiply each of the subscripts of the empirical formula by this number to find the molecular formula.

Example: Find the Empirical Formula

→*What is the empirical formula of a compound with the following percent composition?*

40.00% C, 6.72% H, 53.29% O

Assuming a 100g sample, we have 40.0g of C, 6.72g of H, and 53.3g of O.

Converting all of these to moles, we have 3.33mol of C, 6.67mol of H, and 3.33mol of O.

The smallest of these molar values if 3.33, so we divide each by 3.33 to find the subscripts.

The values for the subscripts are 1, 2, and 1.

Therefore, the empirical formula for this substance is CH_2O.

Example: Find the Molecular Formula

→*Given the same substance above and a mass of 180.g, what is the molecular formula?*

To find the molecular formula, you need the ratio of the empirical formula's mass to the molecular mass. Therefore, you need to compute the molecular molar mass of the empirical formula, which is roughly 30.g.

The ratio is then $\frac{180g}{30g}$, which equals 6.

Now, simply multiply the subscripts on the empirical formula to find the molecular formula. The molecular formula is $C_6H_{12}O_6$.

Average Atomic Mass

The average atomic mass of an element is the molecular molar mass seen on the periodic table. It is a weighted average of all of the element's isotopes.

To find the average atomic mass, you need a list of an element's isotopes, their occurrence in nature as a percent, and a mass for each.

Multiply each isotope's mass by its abundance, and then add the products to find the average atomic mass of the element.

Example: Average Atomic Mass of Silicon

→*Given the table of isotopes, masses, and abundances below, calculate the average atomic mass of silicon.*

Isotope name	Isotope mass (amu)	Relative Abundance
Silicon-28	27.98	92.21
Silicon-29	28.98	4.70
Silicon-30	29.97	3.09

Multiply each mass by its abundance.

$$27.98 \times 0.9221 = 25.80$$

$$28.98 \times 0.0470 = 1.362$$

$$29.97 \times 0.0309 = 0.9261$$

Now, add the products to determine the average atomic mass.

$$25.80 + 1.362 + 0.9261 = 28.09amu$$

The average atomic mass of silicon is 28.09amu.

Stoichiometry

This guide does not cover stoichiometry because it is the most recent topic covered in class.

You have reached the end of this study guide. Good luck on your final.