1 Intro to Chemistry

The first section we covered was an introduction to chemistry, which was mainly a review of things we learned in previous years.

1.1 Scientific Method

A scientific method is a systematic approach used in scientific study, whether it is chemistry, biology, physics, or another science. Ask a question → Do background research → Construct a hypothesis → Experiment → Analyze/Conclude → Communicate results.

1.2 Observations

There are two types of data: qualitative and quantitative.

Qualitative: Information that describes color, odor, shape, or some other physical characteristic (related to 5 senses)
Quantitative: Numerical information (temperature, pressure, volume, quantity)

1.3 Accuracy and Precision

Accuracy: How close you are to the correct answer
Precise: How repeatable your results are.

2 Analyzing Data

This section covers scientific notation, factor label, and significant figures.

2.1 Scientific Notation

Scientific notation is used to express very large or very small numbers in a more condensed, readable method. It is expressed as a number between 0 and 10 multiplied by a power of ten.

\[ 1234 = 1.234 \times 10^3 \]
\[ .0314 = 3.14 \times 10^{-2} \]

2.2 Significant Figures

General rules:
1. All non-zero digits are significant.
2. Zeros to the left of non-zero digits are NEVER significant.
3. Zeros between non-zero digits are ALWAYS significant.
4. Zeros to the right of non-zero digits are significant ONLY if a decimal point is shown.
2.2 Significant Figures (cont.)

<table>
<thead>
<tr>
<th>Number</th>
<th>Sig Figs</th>
</tr>
</thead>
<tbody>
<tr>
<td>.025</td>
<td>2</td>
</tr>
<tr>
<td>408</td>
<td>3</td>
</tr>
<tr>
<td>500</td>
<td>1</td>
</tr>
<tr>
<td>30.0</td>
<td>3</td>
</tr>
<tr>
<td>50,000</td>
<td>5 (note the period!)</td>
</tr>
<tr>
<td>400,002</td>
<td>6</td>
</tr>
<tr>
<td>7.002</td>
<td>4</td>
</tr>
</tbody>
</table>

Multiplying and Dividing: The answer must be rounded off to the same number of significant figures as the least accurate measurement used in the calculation.

34.0 (3 sig figs) x .0921084 (6 sig figs) = 3.1316856
Because the least accurate measurement had 3 sig figs, the answer would be rounded to 3.13.

Adding and Subtracting: The answer must be rounded off to the same column (ones, tenths, hundredths, etc.) as the least precise measurement used in the calculation.

246.58 (hundredths) – 87.3 (tenths) = 159.28
Because the least precise measurement was in the tenths column, the answer would round to 159.3.

Special cases: infinite sig figs.
Counted numbers are numbers with an infinite amount of significant figures.
For example, then number of students in a classroom would be a counted number because it can simply be counted.

2.3 Factor Label

Given
Determine
Convert!

In factor label (dimensional analysis) problems, you are simply using conversion factors to convert something from one unit to another unit.
Remember to write down your given, what you need to determine, and your conversion factors. Being able to see these on paper while doing your problem makes setting up the equation much easier.

If the problem asked you to convert centimeters to kilometers:

\[
\frac{2.56 \text{ cm}}{1} \times \frac{0.01 \text{ m}}{1 \text{ cm}} \times \frac{1 \text{ km}}{1000 \text{ m}} = 0.0000256 \text{ km}
\]
2.3 Factor Label (cont.)

If the problem asked you to convert Joules to Calories:

\[
\left( \frac{551000\text{J}}{1} \right) \left( \frac{1\text{calorie}}{4.184\text{J}} \right) \left( \frac{0.001\text{Cal}}{1\text{calorie}} \right) = 132\text{Cal}
\]

2.4 Percent Error

\[
\%\text{error} = \left( \frac{\text{Actual} - \text{Experimental}}{\text{Actual}} \right) \times 100\%
\]

2.5 Density Calculations

The equation for density is \( D = \frac{m}{V} \), where \( D \) is the density, \( m \) is the mass, and \( V \) is the volume.

Using the fact that \( V = lwh \) and a little bit of algebra, we can manipulate the density equation to solve for length, width, and height.

For example, to find the height:

Remember that water has a density of 1!

3 Matter

In this section, we covered the properties and changes of matter, along with elements, compounds, and mixtures.

3.1 Physical Properties and Changes

A physical property is anything you can measure/notice about a substance without “harming” it. These include color, size, shape, volume, mass, and density.

Physical changes are changes in physical properties. Some examples are water evaporating and grinding beef into a hamburger.

3.2 Chemical Properties and Changes

If you measure a chemical property, you might harm/damage/alter the substance and may not be able to revert it to its original state. Chemical properties include flammability and reactivity.

Some examples of chemical changes include wood burning, silver tarnishing, and food digesting.

There are five signs to look for help to determine if a chemical reaction has occurred:

1. Light is formed
2. Temperature change
3. Color change
4. Gas produced (bubbles) (scent)
5. Precipitate/new substance formed
3.3 Elements, Compounds, Mixtures

Element: A basic substance that cannot be simplified

Pure Substances
Atom: Simplest particle in nature that represents an element
Molecule: Two or more atoms chemically combined
Compound: Two or more atoms chemically combined, atoms **must** be different

- Atom of an element
- Molecule of an element
- Compound

Mixtures
Mixtures are created when elements and/or compounds are mixed together **without chemical reactions** occurring. They retain the individual properties of their component elements and/or compounds. They are not pure substances because the particles that make them up are not identical to one another. They can be separated by physical means. They can be a combination of two or more elements, two or more different compounds, or a combination of elements and compounds.

Homogeneous mixtures (AKA solutions): Made up of individual atoms and/or tiny molecules, components are uniformly distributed throughout the mixture
Example: Air, blood

Heterogeneous mixtures: Made up of larger molecules, components not uniformly distributed
Example: Rocks, soup, pizza

Heterogeneous mixtures can be broken down into colloid**s** and suspensions.

Colloid: Made up of medium sized particles that never settle. They often have a cloud/fuzzy appearance. You can determine if a substance is a colloid by seeing if it displays the **Tyndall Effect**, which is when light scatters when it is shone through a colloid (headlights in fog)
Example: Smoke, fog, whipped cream, milk

Suspension: Made up of larger particles that eventually settle.
Example: Muddy water, paint

Separating Mixtures
Filtration Separates a solid which has not dissolved in a liquid
Solid in remaining filter paper is residue, liquid that passed through filter paper is **filtrate**

Crystallization Separates a solid which has been dissolved in a liquid
Dissolved substance is called **solute**, liquid used for dissolving is called solvent

**Saturated solution** is a solution which has as much solid dissolved in it as it can possibly contain
3.3 Elements, Compounds, Mixtures (cont.)

Separating Mixtures
Distillation Separates one liquid from a mixture of liquids
Collected liquid is called the distillate
Chromatography As the solvent rises up the filter paper, the individual dyes within the colored ink spots are separated and different dyes travel different distances up the paper
The solvent itself travels furthest up the filter paper leaving a line called the solvent front
Most soluble is highest up the paper, least soluble is lowest
Sedimentation/ Insoluble heavy particles in a liquid are allowed to settle down
Decanting Typically used if the insoluble solid has a different density than the liquid
Decantation is the process by which a clear liquid obtained after sedimentation is transferred into another container, without disturbing the settled particles

3.4 Solubility

Solubility curves tell us what mass of solute will dissolve in 100. grams (or 100. mL) of water over a range of temperatures.
Unsaturated Solution: More solute could be dissolved at that temperature
Supersaturated Solution: Holds more solute than it normally could dissolve in that volume of solvent

When solving a factor label problem, think of the graph as a “conversion factor” and the number provided as a “given.” Make sure you know exactly what you want to solve for, and follow the steps that will lead you to that answer.

Test problems could ask:
Saturated, unsaturated, or supersaturated?
How much of this substance will dissolve in x grams of water?
How much of this substance can be dissolved at x°C?
What is the minimum amount of water you would need to dissolve x grams of this substance at y°C?
What would happen if a saturated solution of this substance in x grams of water was cooled from y°C to z°C?

You can find hundreds of practice problems on the internet.

Table C. Solubility Curves

![Solubility Curves Graph]

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4 Atomic theory And Structure

In this section, we learned about atomic theory, different experiments, the structure of atoms, and isotopes.

4.1 Atomic Theory

John Dalton’s Atomic Theory
1. Elements are made of extremely small particles called atoms.
2. Atoms of a given element are identical in size, mass, and other properties; atoms of different elements differ in size, mass, and other properties.
3. Atoms cannot be subdivided, created, or destroyed.
4. Atoms of different elements combine in simple whole-number ratios to form chemical compounds.
5. In chemical reactions, atoms are combined, separated, or rearranged.

Number 2 and 3 have been proven to be incorrect with modern technologies, as isotopes of atoms can be made and it is possible to split an atom.

4.2 Cathode Ray Tube Experiment

In this experiment, J. J. Thomson was able to hypothesize that the light beam formed inside the cathode ray tube was made up of negatively charged particles instead of a previously believed negatively charged light beam. He called these particles electrons.

4.3 Gold Foil Experiment

The diagram of the experiment is on the right.
4.3 Gold Foil Experiment (cont.)

In this experiment, Ernest Rutherford bombarded a thin piece of gold foil with alpha particles. However, while most particles went through unaffected, some particles deflected off of the foil. This led Rutherford to hypothesize every atom had a small, positively charged center, which he called the nucleus.

4.3 Atoms – History

Democritus (400 BC) Proposed that matter is made up of about 100 tiny particles called atoms, but no one believed him
Aristotle (350 BC) Denied the existence of atoms, said matter was made of the four elements (air, earth, fire, water), everyone believed him
400 BC to 1600 AD Dark ages of chemistry, idea of atoms was lost
Robert Boyle (1661) Renews belief in the idea of atoms through experiments with gases
Antoine Lavoisier (1780) “Father of Modern Chemistry,” developed the theory of the Conservation of Matter
John Dalton (1800) Developed atomic theory
William Crookes (1879) Discovered a negatively charged beam in an electrified vacuum tube
JJ Thomson (1897) Proved that Crookes’ beam was really a stream of electrons, realized that electrons must be smaller than the atom, developed “plum pudding” model of atom
Ernest Rutherford (1911) Performed gold foil experiment, won Nobel Prize in chemistry
Niels Bohr (1913) Electrons are in certain shells/energy levels at specific distances from the nucleus, each level has specific amounts of energy, compared electrons orbiting the nucleus of an atom to planets orbiting around the sun
Erwin Schrodinger (1925) Regions where electrons are most likely found are called “electron clouds” or “orbitals”
James Chadwick (1932) Discovered the neutron
1980’s First atoms seen with scanning tunneling microscopes
1995 Fermi lab scientists study and detect six types of quarks, which are particles believed to make up protons and neutrons

4.4 Atoms - Structure

All atoms are composed of a nucleus (protons and neutrons) and electrons revolving around that nucleus. An atom generally consists of three fundamental parts:

- Electron: Negative charge
- Proton: Positive charge, # of protons tells you what atom it is
- Neutron: No charge (neutral)

---------------------- Atomic number. This is how many protons and electrons an element has.

6

C

Carbon 12.0

---------------------- Average atomic mass. This is the average mass of all known isotopes of an element. Also, the number of neutrons is equal to the difference between the atomic number and the average atomic mass.
4.5 Atoms – Isotopes

Isotopes are atoms of the same element with different masses (different number of neutrons). This definition leads to the fact that all atoms are in fact isotopes. It's impossible that there would be an element with a mass of exactly 26.982... (Aluminum); that number is just the average mass of all known isotopes.

A set of isotopes can be written like this:

\[ ^{17}_{8}O \text{ and } ^{16}_{8}O \]

where the mass number is on top, the atomic number is on bottom, and the symbol of the element is on the right.

Note that you can tell that it is a set of isotopes because of the differing mass numbers.

4.6 Atoms – Ions

When an atom is neutral, the number of protons is equal to the number of electrons. If an atom becomes charged (positively or negatively), it is called an ion.

If an ion has a negative charge, it is an anion.

If an ion has a positive charge, it is a cation.

You can determine the charge of an atom by subtracting the number of electrons from the number of protons.

4.7 Atoms – Average Atomic Mass and Percent Abundance

Average Atomic Mass

To calculate:

1. Convert the percentages into decimals (this percentage is known as its relative abundance or percent abundance).
2. Multiply the percentage of each isotope by its respective mass.
3. Add the numbers from step 2 together.

Percent Abundance

1. Because percent abundances will always add up to 100%, assign one isotope to have a percentage of \(x\) and the other isotope to have a percentage of \(1-x\).
2. Look up the average atomic mass of the atom on the periodic table.
3. Set up your problem like this:

\[ x \times (\text{Mass of Isotope A}) + (1 - x) \times (\text{Mass of Isotope B}) = \text{Average Atomic Mass} \]

4. Solve for \(x\). Multiply by 100 to get percent abundance of Isotope A.
5. Subtract that percentage from 100% to get the percent abundance of Isotope B.

5 Electrons

In this section, we learned about light, orbitals, electron configurations, orbital box diagrams, periodic table family names, and periodic trends.
5.1 The Electromagnetic Spectrum

The electromagnetic spectrum shows radiation over a wide range of wavelengths. Light is created when electrons transition from higher to lower energy levels.

As you can see, as elements jump to higher levels, energy is absorbed. When they jump down, energy is released. The color of the light depends on which level the electrons jump down from. The diagram is fittingly called an energy level diagram.

5.2 Atomic Orbitals

Orbitals are different types of electron clouds, or places where you can find electrons.

5.2 Atomic Orbitals (cont.)

We learned about 4 different types of orbitals: s, p, d, and f.
5.2 Atomic Orbitals (cont.)

Pauli Exclusion Principle: Each orbital can only hold up to two electrons
s: 2 electrons
p: 6 electrons
d: 10 electrons
f: 14 electrons

Aufbau Principle: Electrons would like to go into the lowest energy level they can
(low energy) s → p → d → f (high energy)

To determine the maximum number of electrons an energy level can hold, do: \(2n^2\), where \(n\) is 1-7 (period rows)

Electron Configurations
An electron configuration is a description of where all the electrons are in an atom.

Here is the general order of electron configurations:
1s, 2s, 2p, 3s, 3p, 4s, 4p, 5s, 5p, 6s, 6p, 7s, 5f, 6d, and 7p

Electron Configurations

For example, the electron configuration for oxygen would be:
1s\(^2\)2s\(^2\)2p\(^4\)
Iron:
1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)4s\(^2\)3d\(^6\)

To shorten electron configurations, put the highest noble gas that is before the element in front of the configuration.
Iron (shortened):
[Ar]4s\(^2\)3d\(^6\)

Orbitals are the most stable when they are either full or half full. Electrons will sometimes move into a different orbital to help make an atom more stable.
A good example would be the electron configuration for Copper.
Following the normal method, you get:
1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)4s\(^2\)3d\(^9\)
However, in actuality, it’s:
1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)4s\(^1\)3d\(^10\)

5.3 Orbital Diagrams

Orbital diagrams help visualize electron configurations. Each box in an orbital diagram represents an orbital, and each arrow represents an electron.

Hund’s Rule
Electrons will spread out among open orbitals before they pair up.
To put it simply: Draw all your up arrows in the boxes before you start drawing the down arrows.

Sodium:

\[
\begin{array}{cccc}
\uparrow & \uparrow & \uparrow & \uparrow \\
1s & 2s & 2p & 3s \\
\end{array}
\]

Nitrogen:

\[
\begin{array}{cccc}
\uparrow & \uparrow & \uparrow & \uparrow \\
1s & 2s & 2p & \\
\end{array}
\]
5.4 Periodic Table Family Names

<table>
<thead>
<tr>
<th>Group</th>
<th>Family</th>
<th>Valence Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>IA</td>
<td>Alkali Metals</td>
<td>1</td>
</tr>
<tr>
<td>IIA</td>
<td>Alkaline Earth Metals</td>
<td>2</td>
</tr>
<tr>
<td>IIIA</td>
<td>Boron Family</td>
<td>3</td>
</tr>
<tr>
<td>IVA</td>
<td>Carbon Family</td>
<td>4</td>
</tr>
<tr>
<td>VA</td>
<td>Nitrogen Family</td>
<td>5</td>
</tr>
<tr>
<td>VIA</td>
<td>Oxygen Family</td>
<td>6</td>
</tr>
<tr>
<td>VIIA</td>
<td>Halogens</td>
<td>7</td>
</tr>
<tr>
<td>VIIIA</td>
<td>Noble Gases</td>
<td>8</td>
</tr>
</tbody>
</table>

Groups 3-12 are called transition metals.
The last two rows on the periodic table are called lanthanoids and actinoids.

5.5 Periodic Trends

- Ionization energy increases
- Atomic radius decreases
- Electronegativity increases (excluding noble gases)
- Reactivity decreases for metals, increases for nonmetals
- Increasing valence electrons
- Acidity increases

Ionization energy decreases
Atomic radius increases
Electronegativity decreases
Reactivity increases for metals, decreases for nonmetals
Acidity increases

6 Energy

In this section, we learned about endothermic and exothermic reactions and different types of energy problems.
6.1 Exothermic vs. Endothermic

A catalyst speeds up reactions by lowering activation energy.

Heat vs. Temp
Heat: Energy transferred from one object to another because of a temperature difference.
Always goes from hot to cold objects.

Temperature: Average kinetic energy of the particles.
High temperature = moving quickly.
Low temperature = moving slowly.

6.2 $q = mcΔT$

$q$: heat (Joules, KJ, calories)
$m$: mass (g)
$c$: specific heat (heat capacity) = amount of heat you need to get 1 gram of substance to go up by 1°C
$ΔT$: change in temperature ($T_{\text{final}} - T_{\text{initial}}$)

Note: If the question is asking for how many Joules of energy are absorbed, your answer will be positive. If it is asking for how many Joules of energy are released, your answer will be negative.
Just plug your givens into the equation and solve!

Putting a hot iron pan in a sink problem (solving for the final temperature of a ____ and ____ mixture):

$Q_{\text{gained by one object}} = -Q_{\text{lost by another object}}$ OR $m_1c_1ΔT_1 = -m_2c_2ΔT_2$

Don’t forget about the negative!

6.3 Heat of Fusion, Heat of Vaporization

The equation to determine the amount of heat required to convert a solid to a liquid is:

$q = \text{mass} \times ΔH_{\text{fusion}}$
6.3  Heat of Fusion, Heat of Vaporization (cont.)

The equation to determine the amount of heat required to convert a liquid to a gas is:

\[ q = \text{mass} \times \Delta H_{\text{Vaporization}} \]

Note: It takes a lot more heat to evaporate something than it does to melt something.

6.4  Multi-Step Energy Problem

A: “Warming up” a solid (use \(mc\Delta T\))
B: Melting/freezing = (use \(m\Delta H_{\text{fusion}}\))
C: “Warming up” a liquid (use \(mc\Delta T\))
D: Boiling/condensing = (use \(m\Delta H_{\text{vaporization}}\))
E: “Warming up” a gas (use \(mc\Delta T\))

When you solve these, make sure you know what’s happening in each step.
Step A is the energy required to get from the starting temperature to the object’s melting point.
Step B is the energy required to fully melt that object (convert to liquid).
Step C is the energy required to get from a liquid state to its boiling temperature.
Step D is the energy required to convert the object to a gas.
Step E is the energy required to get from a gaseous state to the final temperature.
After you finish solving each of these equations, add up all of your answers to get the final answer.
7 Bonding

In this section, we learned about the properties of ionic and covalent compounds, Lewis dot structures, and the shapes of molecules.

7.1 Ionic Bonds

An ionic bond is made up of a metal and a nonmetal. In an ionic bond, one atom loses an electron and another atom gains one. Metals tend to form cations because they tend to give away electrons, and nonmetals tend to form anions because they tend to gain electrons.

Properties of Ionic Compounds
- Form crystals
- High melting points and high boiling points
- Hard and brittle
- Conduct electricity when dissolved in water

7.2 Covalent Bonds

A covalent bond is made up of nonmetals only. In a covalent bond, electrons are shared between atoms.

To tell if a bond is polar, find the difference between the electronegativities between the atoms. If that number is between 0 and 0.4, you can say it’s a nonpolar covalent bond. If the number is higher than 0.5, you can say it’s a polar covalent bond.

To tell if a molecule is polar, look at the amounts of each element. If there are lots of Carbons and Hydrogens, it’s probably nonpolar. If there are approximately the same number of Carbons and Oxygens, it’s probably polar.

Properties of Covalent Compounds
- Low melting points and boiling points
- Soft and flexible
- Tend to be more flammable than ionic compounds
- Don’t conduct electricity when dissolved in water
- Most don’t dissolve well in water

Note: Network covalent bonds are another type of covalent bond. The elements in this type of bond are mainly from the IVA family and metalloids. They are extremely strong bonds.

7.3 Lewis Dot Structures

Octet rule: When atoms combine to form compounds, they rearrange their electrons to ensure that every atom in the compound is stable and has eight valence electrons. There are some exceptions to this rule – Hydrogen is most stable with 2 valence electrons, Boron and Beryllium can be stable with less than 8 electrons, and Phosphorous and sulfur can be stable with more than 8 electrons.
7.3 Lewis Dot Structures (cont.)

Drawing Lewis Dot Structures for Covalent Compounds
1. Arrange the atoms in the molecule so that they are as symmetrical as possible.
2. Draw each atom’s Lewis Dot structure separately.
3. Show where each atom would bond to another atom by replacing each pair of dots with a line to represent this bond.

\[
\text{H}_2\text{O} \\
\text{H} \cdots \text{O} \cdots \text{H}
\]

The maximum amount of bonds you can have between 2 atoms is 3.

Drawing Lewis Dot Structures for Ionic Compounds
It is customary to show which atom in the compound is donating the electrons and which one is receiving the electrons.

\[
\text{KBr} \\
\text{K}^+ [\cdot \text{Br}^-]
\]

Drawing Lewis Dot Structures for Coordinate Covalent Bonds
Sometimes when elements combine to form compounds, one atom donates both of its electrons into a bond while the other atom doesn’t contribute anything. Coordinate covalent bonding most often occurs in acids and polyatomic ions. It’s typical to draw a circle around the coordinate covalent bonds and draw an arrow starting from the atom donating the electrons to the atom receiving them.

\[
\text{O}_3
\]

Drawing Lewis Dot Structures for Ions
If a group of elements together has a charge, it is called a polyatomic ion. Polyatomic ions bond covalently because they are made up of nonmetals only. If (−) charge, give the electron to the element with the higher electronegativity. If (+) charge, take away the electron from the element with the lower electronegativity.

\[
\text{H}_2\text{O}^+ \\
\left[ \begin{array}{c} \text{H} \\ \cdot \text{O} \cdots \text{H} \\ \text{H} \end{array} \right] +
\]
7.4 Shapes of Molecules

<table>
<thead>
<tr>
<th>Formula</th>
<th>BeCl₂</th>
<th>BCl₃</th>
<th>CH₄</th>
<th>NH₃</th>
<th>H₂O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bonding pairs</td>
<td>2</td>
<td>3</td>
<td>4</td>
<td>3</td>
<td>2</td>
</tr>
<tr>
<td>Valence Electrons</td>
<td>2</td>
<td>3</td>
<td>4</td>
<td>5</td>
<td>6</td>
</tr>
<tr>
<td>Lone Pairs</td>
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<td>0</td>
<td>0</td>
<td>1</td>
<td>2</td>
</tr>
<tr>
<td>Angles between bonding pairs</td>
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<td>120°</td>
<td>109.5°</td>
<td>107°</td>
<td>105°</td>
</tr>
<tr>
<td>Name of shape</td>
<td>Linear</td>
<td>Trigonal Planar</td>
<td>Tetrahedral</td>
<td>Trigonal Pyramid</td>
<td>Bent</td>
</tr>
</tbody>
</table>

8 Moles

In this section, we learned how to calculate molar masses.

8.1 Molar Masses

H₂O:

\[
\frac{2 \text{mole}H}{1} \times \frac{1.0 \text{grams}H}{1 \text{mole}H} = 2.0 \text{grams}H
\]

+ \[
\frac{1 \text{mole}O}{1} \times \frac{16.0 \text{grams}O}{1 \text{mole}O} = 16.0\text{grams}O
\]

Molar mass of H₂O = 18.0 grams/mole