Chapter 1: Chemistry and Measurement

Chemistry (noun): *a science studying the composition, characteristics, properties, and transformations of all matter in the known universe*.

Most modern textbooks start their coverage of chemistry with a discussion of why chemistry is so important to modern life, how it is relevant, why it is the "central science", blah, blah, blah.

Here is the truth; chemistry is important ***to you*** because a group of professionals decided that you must study chemistry in order to earn the degree in your major, or to gain admission to a particular kind of professional school. Whether you agree with their decision or not is irrelevant: you have chosen your major, and your major requires you to study chemistry.

If you don't want to study chemistry, then choose a different major. You don't have to be a chemist, (or a biologist, or a physician, or whatever) in order to be happy and successful. You won't hurt my feelings if you choose to be an artist or a business executive. I would prefer you to be a happy and successful business executive, and not know any chemistry at all, than to have you study a subject you detest. Just don't expect me to justify to you why you should study chemistry; I don't owe you a justification, and neither does science in general nor chemistry in particular.

The above definition of chemistry says it all: all matter, anything you can touch or that can touch you, is within the scope of chemistry. Any type or kind of physical substance you can name or describe is subject to the laws of chemistry.

Normally, the next topic would be a lengthy discussion of the scientific method, hypotheses, theories, laws, and so on. I'm not going to indulge in this waste of paper; instead I'll give you two relatively simple definitions and consider this topic finished.

Theory (noun): *a detailed and elaborate explanation for a set of apparently unrelated facts or observations*.

Scientific law (noun): *a principle or pattern observed in nature, and commonly expressed mathematically*.

An example of a scientific law is the ***law of conservation of mass,*** which tells us that, in a closed system, the total mass is constant. This law is expressed mathematically as:

$$mass_{(initial,total)} = mass_{(final,total)}$$

An example of a theory is ***Atomic theory***, which tells us that all of the different types of matter present in nature are explained as the combination of small particles of individual elements, called ***atoms***. These combinations occur in specific ratios: 1 atom of oxygen for 2 atoms of hydrogen to produce water. Atoms are not permanently changed when they are combined, although the properties of the matter produced by the atoms may be very different from the properties of the original atoms.

Scientific laws and scientific theories are different kinds of knowledge. They are **never** interconverted, and a law is **NOT** a theory that has been "promoted" to a higher level, despite what non-science educated people may believe to the contrary.

Classification of matter

There are three general classes of matter distinguished by their bulk physical properties:

Solid – has a fixed shape, fixed volume, and fixed mass.

Liquid - has no fixed shape, but has fixed volume, and fixed mass.

Gas – has no fixed shape, and no fixed volume, but has fixed mass.

Sometimes the word ***vapor*** is used as a synonym for gas. Sometimes people talk about a fourth phase of matter – ***plasma***. Plasma is an electrically charged gas. While it has many of the properties of ordinary gas, the electric charge gives the plasma some unique properties. Nevertheless, we generally ignore plasmas in introductory chemistry courses.

There are many other ways to organize and classify matter. One important organization is based on the composition of matter.

Element – any substance that can't be broken down by chemical processes into two or more simpler substances. There are currently 114 elements on the periodic table. Each element has a unique **atomic number**, and a unique chemical symbol. The atomic number tells us the number of protons present in the **nucleus** (center) of an atom of the element. The tiniest individual particle of an element, having the same chemical properties as the bulk element, is called an ***atom***. Combining individual atoms of these elements in specific proportions produces all other types of matter.

Compound – a substance composed of two or more elements, chemically combined. For the time being, we won't worry about how atoms "chemically combine". For our purposes, a compound has a fixed and definite ratio of atoms of individual elements, and this ratio is shown in the substances ***chemical formula***. Here are some chemical formulas for ordinary substances:

Water – H_2O Sugar – $C_6H_{12}O_6$

Table salt – NaCl Ammonia – NH_3

In all cases, the chemical formulas list the specific elements making up the substance. The subscripted numbers following the element symbols indicate the number of atoms of that element present in one ***molecule*** of the compound. In water, there are 2 hydrogen atoms and 1 oxygen atom combined together to make a water molecule. In sugar, there are 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms combined together to make a sugar molecule.

A molecule is a particle made up of individual atoms, chemically combined, functioning as a whole, independent unit of matter. A molecule can be made from atoms of one element (not particularly common) or from atoms of different elements (very common).

Compounds follow the ***law of definite proportions*** (also known as the ***law of constant composition***): a pure compound has a fixed and definite ratio, by weight, of individual elements, regardless of the source. Water is 88.81% oxygen by weight, and 11.19% hydrogen by weight. It doesn't matter if the water came from rain, snow, an iceberg, sweat, or from outer space: all water has the same composition by weight.

Elements and compounds are examples of pure matter. If two or more kinds of pure matter are physically combined, then we have a ***mixture*.** There are two general types of mixtures: ***homogeneous*** and ***heterogeneous*.** If a mixture has a single phase, and all substances are uniformly distributed throughout the mixture, then the mixture is called homogeneous. Southern style sweet tea, wine, mayonnaise, white gold, and vanilla ice cream are examples of homogenous mixtures.

If the mixture has two or more phases, and the substances are not uniformly distributed throughout the mixture, then the mixture is heterogeneous. Vegetable soup, Italian salad dressing, rice pudding, and chocolate chip cookies are examples of heterogeneous mixtures.

Figure 1.1 summarizes the classification of matter based on its composition.

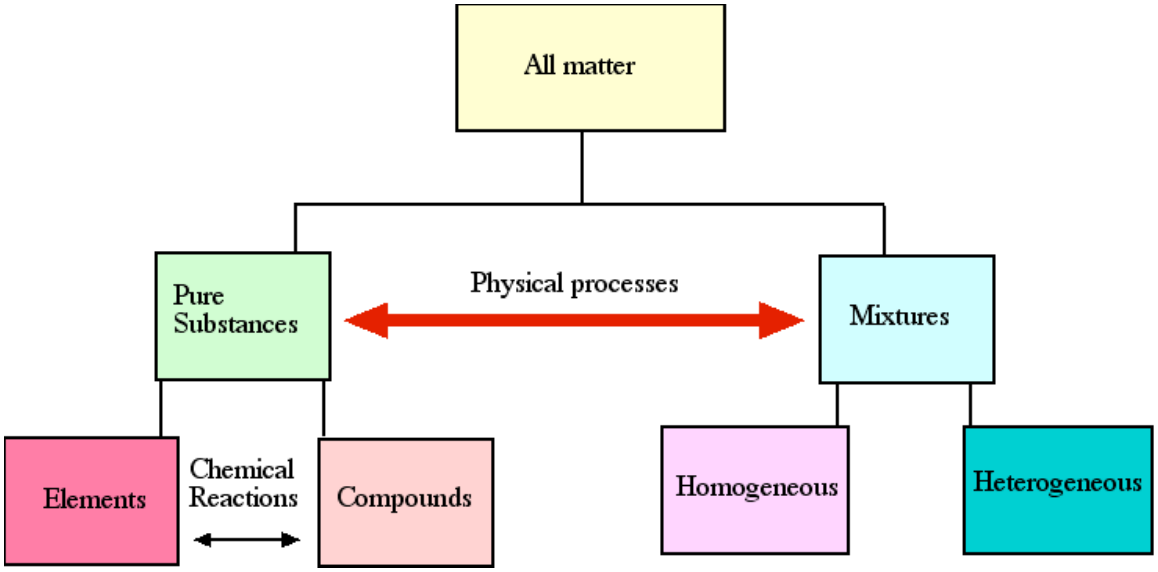

Figure 1.1. Classification of matter based on composition.

Measurements and units

Chemistry is measurement intensive, and the measuring system used by all chemists is the metric system.

It is hardly believable that anyone studying freshman chemistry has not been exposed to the metric system in elementary or high school courses. I am not going to waste time with a long description of the metric system, nor am I going to waste time covering metric-English conversions. You will become comfortable and familiar with the metric system by using the metric system, just as you became comfortable and familiar with the English measuring system by using it.

Scientists use the metric system for two reasons; it is simple and easy to use.

There are seven fundamental units used to measure basic properties. These units are independent from each other, and all other measurement units are derived from these fundamental units. Table 1.1 lists the fundamental units.

Basic Property	Metric Unit
Length	meter (m)
Mass	kilogram (kg)
Time	second (s)
Temperature	kelvin (K)
Electric current	ampere (A)
Amount of substance	mole (mol)
Luminous intensity	candela (cd)

Table 1.1. Fundamental metric units and the properties they measure.

You probably won't use electric current and luminous intensity in general chemistry, and will only rarely use time. The other four properties (length, mass, temperature, amount of substance) are used regularly.

Every other unit is derived from the base units and appropriate mathematics. One unit that is commonly used, but isn't fundamental, is a unit for volume. Consider the box shown below (Figure 1.2.). If we want to find the volume of this box, we multiply the width (W) times the length (L) times the height (H) to get the volume: V = L x W x H. The volume unit would be cubic units of distance: if the length, width, and height were measured in meters, then the volume unit would be meter x meter x meter = meter^3. Volume is a *derived unit*, and you will encounter many other derived units as you progress through this course.

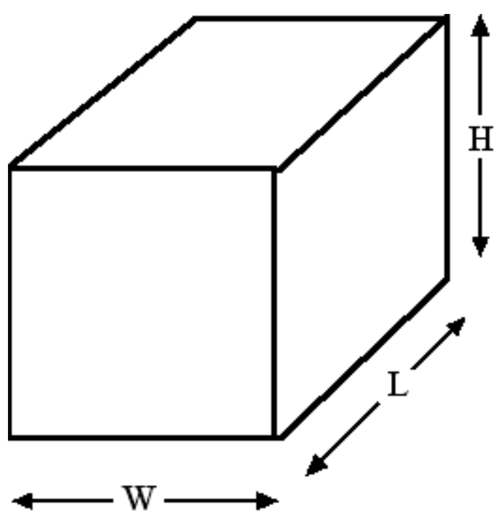

Figure 1.2. A box.

In general, volumes of m^3 are inconveniently large. Instead we use a smaller volume unit, the liter (L). One liter is equal to 1000 cm^3 (cubic centimeter) or 1000 mL (milliliter).

All measured quantities must have the unit included with the measured value. Saying that something is "3.47" is not informative, but saying that something is "3.47 meters" gives me a great deal of information. The unit tells me that the basic property being measured is length, that the metric system is being used, and that the measurement was made to the nearest centimeter (1 cm = 0.01m). Most students develop the habit of dropping off units, and in response most teachers develop the habit of subtracting points from test answers that don't have units.

When calculations are made using base metric units, we commonly derive units that are simple combinations of the base units. For example, if we want a unit to express the idea of *speed* (the distance traveled per unit of time), the unit is meters/second (m/s), where the diagonal line indicates a division. An object traveling 35 meters in 12 seconds is travelling 35/12 = 2.92 m/s.

You might also see this unit in a more "traditional" fraction format such as $\frac{m}{s}$.

Occasionally, a "simpler" metric unit is substituted for a combination of base units. The metric unit for energy is the joule (J), with 1 J equal to 1 $kg\text{-}m^2/s^2$. (The dash "-" in this unit does NOT indicate subtraction. It simply separates one metric unit from another for clarity.)

All measuring systems must be able to handle values ranging from the very small to the very large. The metric system accomplishes this using a set of constant value prefixes to indicate how the units have been increased/decreased in size. Table 1.2 lists these prefixes and their affects on the unit.

Notice that the abbreviation for the prefix is typically one letter (deka and micro are exceptions). Upper case letters are typically used for large units, while lower case letters are used for small units (kilo, hecto, and deka are exceptions). You must use the proper case for the proper unit: zeta (Z) can be easily misunderstood for zepto (z) if you aren't careful! The three most important prefixes are given in bold, and you should definitely learn these.

Prefix name	Affect on unit
yotta (Y)	unit x 10^{24}
zetta (Z)	unit x 10^{21}
exa (E)	unit x 10^{18}
peta (P)	unit x 10^{15}
tera (T)	unit x 10^{12} (1,000,000,000,000)
giga (G)	unit x 10^{9} (1,000,000,000)
mega (M)	unit x 10^{6} (1,000,000)
kilo (k)	**unit x 10^{3} (1000)**
hecto (h)	unit x 10^{2} (100)
deka (deca) (da)	unit x 10
deci (d)	unit ÷ 10
centi (c)	**unit ÷ 100**
milli (m)	**unit ÷ 10^{3} (1000)**
micro (μ)	unit ÷ 10^{6} (1,000,000)
nano (n)	unit ÷ 10^{9} (1,000,000,000)
pico (p)	unit ÷ 10^{12} (1,000,000,000,000)
femto (f)	unit ÷10^{15}
atto (a)	unit ÷10^{18}
zepto (z)	unit ÷10^{21}
yocto (y)	unit ÷10^{24}

Table 1.2. Metric prefixes and the affect on the unit.

Significant figures

The proper use of significant figures is always frustrating for students, so I have adopted a very simple approach: for all calculated values, show 3 non-zero digits (or 3 digits total if one is a "trapped" zero). Consider the following examples:

A. Your calculator shows a value of 19.35467669. Report this as 19.4.

B. A value of 527,988,365,216,363 should be reported as 5.28 x 10^{14}.

C. You calculate an answer of 0.309427867; report this as 0.309. In this case, we have a "trapped" zero separating the 3 and the 9.

D. Your calculator gives a value of 0.000000000224; report this as 2.24×10^{-10}.

E. You have a value of 100,000,000. This should be reported as 1.00×10^{8}.

The last example is a special case – just go with it.

I am not saying that the proper use of significant figures isn't important. It is very important in scientific and engineering calculations. However, it is not one of the most important things for you to learn in general chemistry. If I have to choose between fooling around for one lecture with significant figures, or spend that time teaching something more important, I know what I will choose.

Dimensional analysis and conversion factors

Dimensional analysis is a calculation procedure allowing us to express calculated values with proper units. You have already seen an example of how new units are produced in the previous section. Basically, units follow the same rules of arithmetic that numbers follow.

For addition and subtraction, the units of the quantities being added or subtracted **MUST BE IDENTICAL**. Subtracting 4 apples from 9 apples gives us 5 apples:

$$9 \text{ apples} - 4 \text{ apples} = 5 \text{ apples}$$

Subtracting 4 apples from 9 oranges is a meaningless operation.

Sometimes we will be given values having identical units; other times we will have to change the units of one or more values before we can add or subtract. We can change units using ***conversion factors***. A conversion factor is the ratio of two equivalent quantities. For example, 1 kilogram is equal to 1000 grams. Whether we measure mass in kilograms, or in grams, we can change from grams to kilograms or kilograms to grams using the proper conversion factor.

From any pair of exactly equivalent quantities, two conversion factors can be written: 1-kilogram/1000 grams or 1000 grams/1-kilogram. Both of these conversion factors are equally correct. Which conversion factor do we use in a problem? The choice depends on the units of the quantity we want to convert. If I want to convert 7 kilograms into grams, the conversion factor I use is 1000 grams/1 kilogram. I perform the following calculation:

$$7 \text{ kilograms} \times \frac{1000 \text{ grams}}{1 \text{ kilogram}} = 7000 \text{ grams}$$

This example shows the most important idea in using conversion factors: select the conversion factor that has the desired units on top. Then, when you multiply your starting quantity by the conversion factor, the desired units will carry through the calculation.

Sometimes, several conversion factors must be used before the final result is determined. For example: I have 5.0 kilograms of iron. If I want to express this amount of iron in milligrams, I need to use two conversion factors: 1000 grams/1 kilogram and 1000-milligrams/1 gram. I then perform the following calculation:

$$5.0 \text{ kilograms x } \frac{1000 \text{ grams}}{1 \text{ kilogram}} \text{ x } \frac{1000 \text{ milligrams}}{1 \text{ gram}} = 5{,}000{,}000 \text{ milligrams } (5.00 \text{ x } 10^6 \text{ mg})$$

For multiplication and division, units multiply together or divide each other. We have seen that units of length are multiplied together in calculating volume. Units that aren't the same also multiply together; under some circumstances we can have units of kilograms multiplied by meters to get kilogram-meter (kg x m = kg-m). Notice again that the dash is a separator for clarity; it does NOT indicate that meters are being subtracted from kilograms. In division, units are written using a diagonal slash showing a ratio of units. Speed has units of meter/second (m/s) and is read as "meters per second". The joule (J) can also be written as kg-m^2/s^2, which is read as "kilogram-meter squared per second squared".

If identical units are divided, then the units disappear; they are said to "cancel out". This is exactly what happens with our conversion factors. Consider the earlier example of converting 7 kilograms into grams. The conversion factor we choose is 1000 grams/1 kilogram. The detailed calculation is shown below:

$$7 \text{ kilograms} \times \frac{1000 \text{ grams}}{1 \text{ kilogram}} = \frac{7000 \text{ kilogram-grams}}{1 \text{ kilogram}} = 7000 \text{ grams}$$

Notice that the intermediate unit is kilogram-grams/kilogram. Units of kilogram and gram were multiplied together to create a new unit (kilogram-grams), and then the new unit was divided by kilogram to produce grams.

Sometimes, division produces a pure, dimensionless number. These situations are rare and occur when the units for numerator and denominator are identical.

One special pair of conversions involves temperature scales. The metric temperature scale is Kelvin, and Kelvin temperatures are given as "______ Kelvin"; room temperature is 298 Kelvin. Strictly speaking, we don't us a degree symbol – we don't write 298 ºK and we don't say "two hundred and ninety eight degrees Kelvin". However, if you should include the degree symbol, it's not really that big of a deal for now.

While Kelvin is the metric temperature scale, we almost never measure temperatures directly in Kelvin. Instead, we measure temperature using the Celcius temperature scale (ºC) and convert to Kelvin whenever necessary. The conversion is:

$$K = {}^{\circ}C + 273.15$$

Frequently, the decimal "0.15" is omitted, and K = ºC + 273 is used.

Another temperature conversion frequently encountered is converting from Celcius to Fahrenheit (ºF). This conversion is:

$$\left({}^{\circ}C \times 1.8\right) + 32 = {}^{\circ}F$$

Water boils at 100 ºC. The equivalent Fahrenheit temperature is:

$$(100\ {}^{\circ}C \times 1.8) + 32 = (180) + 32 = 212\ {}^{\circ}F$$

Other conversions will be introduced when necessary. For now, worry about metric conversion factors and temperature conversions.

Practice all of the homework problems over and over again, until you can perform the calculations rapidly and accurately.

Vocabulary. The following terms are defined and explained in the text. Make sure that you are familiar with the meanings of the terms as used in chemistry. Understand that you may have been given incomplete or mistaken meanings for these terms in earlier courses. The meanings given in the text are correct and proper.

Chemistry	**Theory**	**Scientific law**
Conversion factor	**Atom**	**Solid**
Liquid	**Gas**	**Vapor**
Plasma	**Element**	**Atomic number**
Nucleus	**Compound**	**Chemical formula**
Molecule	**Mixture**	**Homogeneous**
Heterogeneous	**Law of conservation of mass**	**Law of definite proportions (Law of constant composition)**

Homework: Use the dimensional analysis procedure to make the indicated calculations. Practice these problems over and over, until solving them is fast and correct.

1. How many grams of iron are in 3.75 kg of iron?

2. What is the distance between Nashville and Murfreesboro centimeters? The distance between Nashville and Murfreesboro is 48 km. Equivalent quantities: 1 meter = 100 centimeters; 1 kilometer = 1000 meters.

3. The density of mercury is 13.5 g/mL. What is the density of mercury in grams/cm^3? 1 cm^3 = 1 mL.

4. The wavelength of green light is 600 nm (nanometers). What is the wavelength of green light in meters? 1 meter = 1,000,000,000 nm.

5. How many milliliters of water are in 15 liters of water? Equivalent quantities: 1 liter = 1000 milliliters.

6. The density of gold is 19.3 g/cm^3. What volume does 500 grams of gold occupy?

7. Complete the following table by filling in appropriate values.

17.5 g	mg	kg
cm	0.339 m	mm
86,400 s	ks	Ms
fg	μg	37mg
570 mL	L(liters)	cL
57 K	mK	kK
mol	mmol	0.58 kmol
nm	55.9 cm	m
ks	ms	245 s
mg	5.99 kg	g

8. Density is the ratio of mass to volume, and is calculated from $density = \frac{mass}{volume}$. Complete the following table.

15.01 g	37.9 mL	g/mL
kg	19.0 cm^3	2.95 g/cm^3
21.5 g	mL	8.34 g/mL
0.995kg	1510 cm^3	g/cm^3
g	15.03 mL	1.95 g/mL
2100 g	m^3	133 kg/m^3
g	27.6 μL	15.3 mg/L
1175 mg	15 dL	g/L
g	1 m^3	1.225 kg/m^3
1.00 g	mL	1.116 g/mL

9. Make the indicated conversions between temperature scales.

158 K	^oC	^oF
K	77 ^oC	^oF
K	^oC	350 ^oF
25 K	^oC	^oF
K	212 ^oC	^oF
K	^oC	98 ^oF
610 K	^oC	^oF
K	55 ^oC	^oF
K	^oC	1100 ^oF
1200 K	^oC	^oF

10. State whether each of the following represents an atom or a molecule.

Au	CCl_4	H_2O	Na	Pb
NH_3	KCl	Xe	NaCl	H_2SO_4
Zn	N	Fl	CH_3OH	$Zn(OH)_2$
O_2	Fe	C_6H_6	NaH_2PO_4	Ca

11. For the following compounds, indicate which elements are present (by symbol, not name) in the compound and indicate how many atoms of each element are present in one molecule. For example: $NaNO_3$ has 1 Na, 1 N, and 3 O.

a. NaH_2PO_4

b. $Zn(OH)_2$

c. $Ni(NO_3)_2$

d. CCl_4

e. NH_3

f. H_2O

g. $NaCl$

h. H_2SO_4

i. CH_3OH

j. KCl

k. C_6H_6

Answers:

1. $3.75\text{ kg} \times \dfrac{1000\text{ g}}{1\text{ kg}} = 3750\text{ g (or } 3.75 \text{ x } 10^3\text{ g)}$

2. $48\text{ km x } \dfrac{1000\text{ m}}{1\text{ km}} \text{ x } \dfrac{100\text{ cm}}{1\text{ m}} = 4{,}800{,}000\text{ cm (or } 4.80 \text{ x } 10^6\text{m)}$

3. $13.5\ \dfrac{\text{g}}{\text{mL}} \text{ x } 1\ \dfrac{\text{mL}}{\text{cm}^3} = 13.5\ \dfrac{\text{g}}{\text{cm}^3}$

4. $600\text{ nm x } \dfrac{1\text{ meter}}{1{,}000{,}000{,}000\text{ nm}} = 6.00 \text{ x } 10^{-7}\text{m}$

5. $15\text{ L x } \dfrac{1000\text{ mL}}{\text{L}} = 15{,}000\text{ mL (or } 1.5 \text{ x } 10^4\text{ mL)}$

6. $\dfrac{500\text{ g}}{19.3\text{ g/cm}^3} = 25.9\text{ cm}^3$

7.

17.5 g	**17,500 mg**	**0.0175 kg**
33.9 cm	0.339 m	**339 mm**
86,400 s	**86.4 ks**	**0.0864 Ms**
3.7 x 10^{13} fg	**37,000 µg**	37 mg
570 mL	**0.570 L**	**57.0 cL**
57 K	**57,000 mK**	**0.057 kK**
580 mol	**580,000 mmol**	0.58 kmol
5.59 x 10^8 nm	55.9 cm	**0.559 m**
0.245 ks	**245,000 ms**	245 s
5,990,000 mg	5.99 kg	**5,990 g**

8.

15.01 g	37.9 mL	**0.396 g/mL**
0.05605 kg	19.0 cm^3	2.95 g/cm^3
21.5 g	**2.58 mL**	8.34 g/mL
0.995kg	1510 cm^3	**0.659 g/cm^3**
29.3 g	15.03 mL	1.95 g/mL
2100 g	**0.0158 m^3**	133 kg/m^3
4.22 x 10^{-7} g	27.6 μL	15.3 mg/L
1175 mg	15 dL	**0.783 g/L**
1225 g	1 m^3	1.225 kg/m^3
1.00 g	**0.896 mL**	1.116 g/mL

9.

158 K	-115 ^oC	-175 ^oF
350 K	77 ^oC	171 ^oF
450 K	177 ^oC	350 ^oF
25 K	-248 ^oC	-414 ^oF
485 K	212 ^oC	414 ^oF
310 K	37 ^oC	98 ^oF
610 K	337 ^oC	639 ^oF
328 K	55 ^oC	131 ^oF
866 K	593 ^oC	1100 ^oF
1200 K	927 ^oC	1701 ^oF

10.

Au **Atom**	CCl_4 **Molecule**	H_2O **Molecule**	Na **Atom**	Pb **Atom**
NH_3 **Molecule**	KCl **Molecule**	Xe **Atom**	NaCl **Molecule**	H_2SO_4 **Molecule**
Zn **Atom**	N **Atom**	Fl **Atom**	CH_3OH **Molecule**	$Zn(OH)_2$ **Molecule**
O_2 **Molecule**	Fe **Atom**	C_6H_6 **Molecule**	NaH_2PO_4 **Molecule**	Ca **Atom**

11.

a. NaH_2PO_4 **1 Na, 2 H, 1 P, 4 O**

b. $Zn(OH)_2$ **1 Zn, 2 O, 2 H**

c. $Ni(NO_3)_2$ **1 Ni, 2 N, 6 O**

d. CCl_4 **1 C, 4 Cl**

e. NH_3 **1 N, 3 H**

f. H_2O **2 H, 1 O**

g. NaCl **1 Na, 1 Cl**

h. H_2SO_4 **2 H, 1 S, 4 O**

i. CH_3OH **1 C, 4 H, 1 O**

j. KCl **1 K, 1 Cl**

k. C_6H_6 **6 C, 6 H**