Lesson 1: The Study of Chemistry

- ◯ Describe how chemistry and engineering helped transform metal into an inexpensive structural material.
- ◯ Explain the usefulness of the macroscopic, microscopic, and symbolic perspectives in understanding chemical systems.
- ◯ Express the results of calculations using the correct number of significant figures.

Definition of Terms

<u>Matter</u> is anything that has mass and can be observed. (Q: Is light a matter?)

<u>Atoms</u> are unimaginably small particles that cannot be made any smaller and still behave like a chemical system.

<u>Molecules</u> are groups of atoms held together so that they form a unit whose identity is distinguish different from the atoms alone.

<u>Element</u> is a substance that cannot be separated into simpler substances by chemical
means.

<u>Compound</u>, a substance composed of atoms of two or more elements chemically united in fixed proportions.

<u>Mixture</u> is a combination of two or more substances.

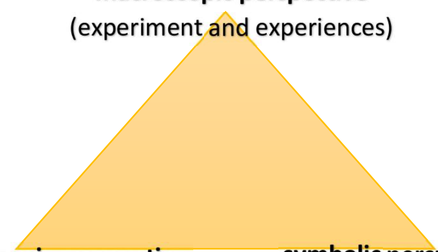
<u>Homogeneous Mixture</u> is a mixture in which the composition is uniform. Also called a solution.

<u>Heterogeneous Mixture</u> is a mixture in which the composition varies.

Chemistry: A Science for Twenty-First Century

- Chemistry is the study of matter and the changes it undergoes.
- Chemistry is the science of the composition and structure of materials and of the changes that materials undergo.
- Chemistry is often called the "central science", because a basic knowledge of chemistry is essential for students of biology, physics, geology, ecology, and many other subjects.
- Chemistry is the science of matter, and since all engineering designs involve matter, the links between chemistry and engineering are many.
- Chemistry is an empirical science. It relies on experimental observations to develop an understanding of matter.

Three levels of understanding or perspective on the nature of chemistry:



- The Macroscopic Perspective is the viewpoint of chemistry focusing on samples of matter that are large enough to be seen, measured, or handled easily.
- Microscopic Perspective is the viewpoint of chemistry focusing on samples of matter at the atomic and molecular level, where samples cannot be seen, measured, or handled easily. Note that this scale is smaller than the resolution of a traditional microscope. Also called the particulate perspective.
- Symbolic Perspective is the viewpoint of chemistry focusing on symbolic representations of the substances involved through formulas, equations, etc.

Classifications of Matter:

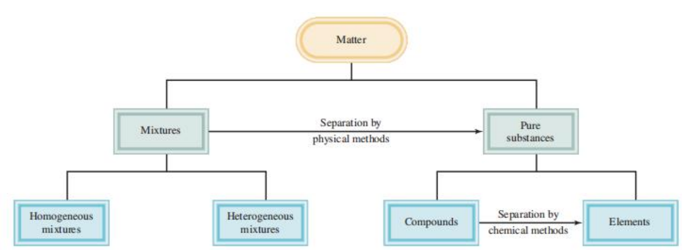


Figure 1.5 *Classification of matter.*

Matter is anything that occupies space and has mass. Matter can be classified as:

1. <u>Substance</u>- is a form of matter that has a definite (constant) composition and distinct properties. Substance can be:
   - <u>element</u> is a substance that cannot be separated into simpler substances by chemical means.
   - <u>substance</u> composed of atoms of two or more elements chemically united in fixed proportions.
2. <u>Mixture</u>- is a combination of two or more substances in which the substances retain their distinct identities. Mixtures can be :

- homogeneous mixture in which the composition of the mixture is the same throughout.
- heterogeneous mixture because the composition is not uniform.

Basic Properties of Matter

Substances are identified by their properties as well as by their composition. All properties of matter are either extensive or intensive and either physical or chemical. Extensive properties, such as mass and volume, depend on the amount of matter that is being measured. Intensive properties, such as density and color, do not depend on the amount of matter. Both extensive and intensive properties are physical properties, which means they can be measured without changing the substance's chemical identity..

1. Physical property - Property that can be observed or measured while the substance being observed retains its composition and identity.
   - Extensive Property – vary with the amount of the substance
     - Mass (m)
     - Volume (V)
   - Intensive Property - do not depend on the amount of the substance
     - Color
     - Density (ρ or D)
     - Melting point/boiling point
     - Electrical conductivity

2. Chemical property - Property of a substance that is associated with the types of chemical changes that the substance undergoes.

- Heat of combustion is the energy released when a compound undergoes complete combustion (burning) with oxygen. The symbol for the heat of combustion is ΔHc.
- Chemical stability refers to whether a compound will react with water or air (chemically stable substances will not react). Hydrolysis and oxidation are two such reactions and are both chemical changes.
- Flammability refers to whether a compound will burn when exposed to flame. Again, burning is a chemical reaction—commonly a high-temperature reaction in the presence of oxygen.
- The preferred oxidation state is the lowest-energy oxidation state that a metal will undergo reactions in order to achieve (if another element is present to accept or donate electrons).

Three States of Matter:

Three States of Matter:

State	Macroscopic	Microscopic	Microscopic Visualization
SOLID	hard and do not change their shapes easily	molecules are held close together in an orderly fashion with little freedom of motion.	
LIQUID	adapt to the shape of the container in which they are held	close together but are not held so rigidly in position and can move past one another.	
GAS	expands to occupy the entire volume of its container	molecules are separated by distances that are large compared with the size of the molecules.	

The three states of matter can be inter-converted without changing the composition of the substance. Upon heating, a solid (for example, ice) will melt to form a liquid (water). (The temperature at which this transition occurs is called the melting point.) Further heating will convert the liquid into a gas. (This conversion takes place at the boiling point of the liquid.) On the other hand, cooling a gas will cause it to condense into a liquid. When the liquid is cooled further, it will freeze into the solid form.

Measurement

The measurements chemists make are often used in calculations to obtain other related quantities. Different instruments enable us to measure a substance's properties: The meterstick measures length or scale; the buret, the pipet, the graduated cylinder, and the volumetric fl ask measure volume; the balance measures mass; the thermometer measures temperature. These instruments provide measurements of macroscopic properties, which can be determined directly. Microscopic properties, on the atomic or molecular scale, must be determined by an indirect method

SI Units

In 1960, the General Conference of Weights and Measures, the international authority on units, proposed a revised metric system called the International System of Units (abbreviated SI, from the French Système Internationale d'Unites).

Measurements that we will utilize frequently in our study of chemistry include time, mass, volume, density, and temperature.

TABLE 1.2 SI Base Units

Base Quantity	Name of Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electrical current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

TABLE 1.3 Prefixes Used with SI Units

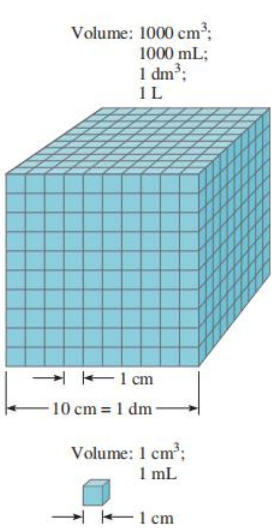
Prefix	Symbol	Meaning	Example
tera-	T	1,000,000,000,000, or 10^{12}	1 terameter (Tm) = 1×10^{12} m
giga-	G	1,000,000,000, or 10^{9}	1 gigameter (Gm) = 1×10^{9} m
mega-	M	1,000,000, or 10^{6}	1 megameter (Mm) = 1×10^{6} m
kilo-	k	1,000, or 10^{3}	1 kilometer (km) = 1×10^{3} m
deci-	d	1/10, or 10^{-1}	1 decimeter (dm) = 0.1 m
centi-	c	1/100, or 10^{-2}	1 centimeter (cm) = 0.01 m
milli-	m	1/1,000, or 10^{-3}	1 millimeter (mm) = 0.001 m
micro-	μ	1/1,000,000, or 10^{-6}	1 micrometer (μm) = 1×10^{-6} m
nano-	n	1/1,000,000,000, or 10^{-9}	1 nanometer (nm) = 1×10^{-9} m
pico-	p	1/1,000,000,000,000, or 10^{-12}	1 picometer (pm) = 1×10^{-12} m

1. Mass and Weight

- Mass is a measure of the amount of matter in an object
- Weight is the force that gravity exerts on an object
- The SI unit of mass is the kilogram (kg). Unlike the units of length and time, which are based on natural processes that can be repeated by scientists anywhere, the kilogram is defined in terms of a particular object.

1 kg = 1000 g = 1 × 10^3 g

2. Volume



- The SI unit of length is the meter (m), and the SI-derived unit for volume is the cubic meter (m3).

$$1\ \text{cm}^3 = (1 \times 10^{-2}\ \text{m})^3 = 1 \times 10^{-6}\ \text{m}^3$$
$$1\ \text{dm}^3 = (1 \times 10^{-1}\ \text{m})^3 = 1 \times 10^{-3}\ \text{m}^3$$

- Another common unit of volume is the liter (L). A liter is the volume occupied by one cubic decimeter. One liter of volume is equal to 1000 milliliters (mL) or 1000 cm3:

$$\begin{aligned} 1\ \text{L} &= 1000\ \text{mL} \\ &= 1000\ \text{cm}^3 \\ &= 1\ \text{dm}^3 \end{aligned}$$

- and one milliliter is equal to one cubic centimeter:

$$1\ \text{mL} = 1\ \text{cm}^3$$

3. Density

- The equation for density is *d = m/V,* where d, m, and V denote density, mass, and volume, respectively. Because density is an intensive property and does not depend on the quantity of mass present, for a given substance the ratio of mass to volume always remains the same; in other words, V increases as m does. Density usually decreases with temperature.
- The SI-derived unit for density is the kilogram per cubic meter (kg/m3).

$$1\ \text{g/cm}^3 = 1\ \text{g/mL} = 1000\ \text{kg/m}^3$$
$$1\ \text{g/L} = 0.001\ \text{g/mL}$$

Sample Problem:

Gold is a precious metal that is chemically unreactive. It is used mainly in jewelry, dentistry, and electronic devices. A piece of gold ingot with a mass of 301 g has a volume of 15.6 cm3. Calculate the density of gold.

$$\begin{aligned} d &= \frac{m}{V} \\ &= \frac{301\ \text{g}}{15.6\ \text{cm}^3} \\ &= 19.3\ \text{g/cm}^3 \end{aligned}$$

Sample Problem:

The density of mercury, the only metal that is a liquid at room temperature, is 13.6 g/mL. Calculate the mass of 5.50 mL of the liquid.

$$\begin{aligned} m &= d \times V \\ &= 13.6 \frac{\text{g}}{\cancel{\text{mL}}} \times 5.50\ \cancel{\text{mL}} \\ &= 74.8\ \text{g} \end{aligned}$$

4. Temperature Scales

- Three temperature scales are currently in use:

1. <u>°F (degrees Fahrenheit)</u>- most commonly used scale in the United States outside the laboratory, defines the normal freezing and boiling points of water to be exactly 32°F and 212°F, respectively.

2. <u>°C (degrees Celsius)</u>- divides the range between the freezing point (0°C) and boiling point (100°C) of water into 100 degrees.

3. <u>K (kelvin)</u>- is the SI base unit of temperature: it is the absolute temperature scale. By absolute we mean that the zero on the Kelvin scale, denoted by 0 K, is the lowest temperature that can be attained theoretically

- To convert degrees Fahrenheit to degrees Celsius:

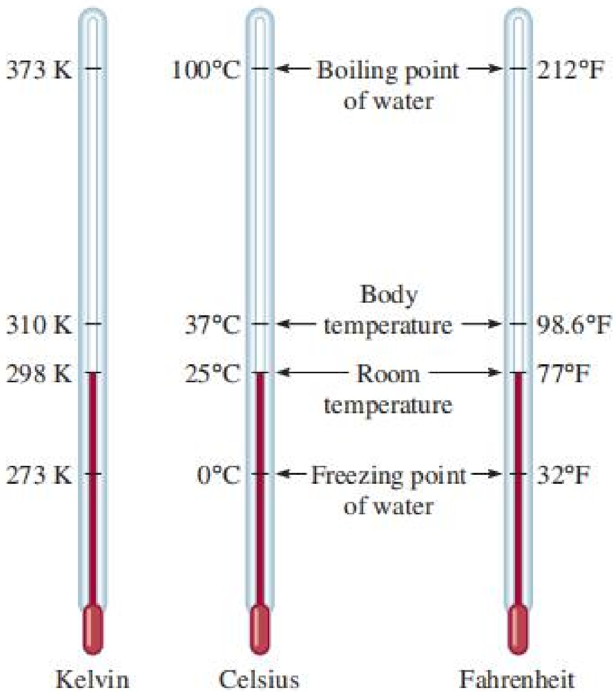
$$?°\text{C} = (°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}}$$

- To convert degrees Celsius to degrees Fahrenheit:

$$?°F = \frac{9°F}{5°C} \times (°C) + 32°F$$

- To convert degrees Celsius to kelvin:

$$?\ K = (°C + 273.15°C)\frac{1\ K}{1°C}$$



Sample Problem:

(a) Solder is an alloy made of tin and lead that is used in electronic circuits. A certain solder has a melting point of 224°C. What is its melting point in degree Fahrenheit?

$$\frac{9°F}{5°C} \times (224°C) + 32°F = \boxed{435°F}$$

(b) Helium has the lowest boiling point of all the elements at 2452°F. Convert this temperature to degrees Celsius.

$$(-452°F - 32°F) \times \frac{5°C}{9°F} = \boxed{-269°C}$$

(c) Mercury, the only metal that exists as a liquid at room temperature, melts at 238.9°C. Convert its melting point to kelvins.

$$(-38.9°C + 273.15°C) \times \frac{1\ K}{1°C} = \boxed{234.3\ K}$$

Scientific Notation

When working with very large and very small numbers, we use a system called scientific notation. Regardless of their magnitude, all numbers can be expressed in the form

N × 10n

Where N is a number between 1 and 10 and n, the exponent, is a positive or negative integer (whole number). Any number expressed in this way is said to be written in scientific notation.

(1) Express 568.762 in scientific notation:

$$568.762 = 5.68762 \times 10^2$$

(2) Express 0.00000772 in scientific notation:

$$0.00000772 = 7.72 \times 10^{-6}$$

Addition and Subtraction

To add or subtract using scientific notation, we first write each quantity—say N1 and N2—with the same exponent n. Then we combine N1 and N2; the exponents remain the same. Consider the following examples:

$$\begin{aligned}(7.4 \times 10^3) + (2.1 \times 10^3) &= 9.5 \times 10^3\\ (4.31 \times 10^4) + (3.9 \times 10^3) &= (4.31 \times 10^4) + (0.39 \times 10^4)\\ &= 4.70 \times 10^4\\ (2.22 \times 10^{-2}) - (4.10 \times 10^{-3}) &= (2.22 \times 10^{-2}) - (0.41 \times 10^{-2})\\ &= 1.81 \times 10^{-2}\end{aligned}$$

Multiplication and Division

To multiply numbers expressed in scientific notation, we multiply N1 and N2 in the usual way, but add the exponents together. To divide using scientific notation, we divide N1 and N2 as usual and subtract the exponents. The following examples show how these operations are performed:

$$\begin{aligned}(8.0 \times 10^4) \times (5.0 \times 10^2) &= (8.0 \times 5.0)(10^{4+2})\\ &= 40 \times 10^6\\ &= 4.0 \times 10^7\\ (4.0 \times 10^{-5}) \times (7.0 \times 10^3) &= (4.0 \times 7.0)(10^{-5+3})\\ &= 28 \times 10^{-2}\\ &= 2.8 \times 10^{-1}\\ \frac{6.9 \times 10^7}{3.0 \times 10^{-5}} &= \frac{6.9}{3.0} \times 10^{7-(-5)}\\ &= 2.3 \times 10^{12}\\ \frac{8.5 \times 10^4}{5.0 \times 10^9} &= \frac{8.5}{5.0} \times 10^{4-9}\\ &= 1.7 \times 10^{-5}\end{aligned}$$

Significant Figures

- It is important to indicate the margin of error in a measurement by clearly indicating the number of significant figures, which are the meaningful digits in a measured or calculated quantity.

- When significant figures are used, the last digit is understood to be uncertain.

Guidelines for Using Significant Figures

We must always be careful in scientific work to write the proper number of significant figures. In general, it is fairly easy to determine how many significant figures a number has by following these rules:

1. Any digit that is not zero is significant. Thus, 845 cm has three significant figures, 1.234 kg has four significant figures, and so on.

2. Zeros between nonzero digits are significant. Thus, 606 m contains three significant figures, 40,501 kg contains five significant figures, and so on.

3. Zeros to the left of the first nonzero digit are not significant. Their purpose is to indicate the placement of the decimal point. For example, 0.08 L contains one significant figure, 0.0000349 g contains three significant figures, and so on.

4. If a number is greater than 1, then all the zeros written to the right of the decimal point count as significant figures. Thus, 2.0 mg has two significant figures, 40.062 mL has five significant figures, and 3.040 dm has four significant figures. If a number is less than 1, then only the zeros that are at the end of the number and the zeros that are between nonzero digits are significant. This means that 0.090 kg has two significant figures, 0.3005 L has four significant figures, 0.00420 min has three significant figures, and so on.

5. For numbers that do not contain decimal points, the trailing zeros (that is, zeros after the last nonzero digit) may or may not be significant. Thus, 400 cm may have one significant figure (the digit 4), two significant figures (40), or three significant figures (400). We cannot know which is correct without more information. By using scientific notation, however, we avoid this ambiguity. In this particular case, we can express the number 400 as 4 × 10^2 for one significant figure, 4.0 × 10^2 for two significant figures, or 4.00 × 10^2 for three significant figures.

A second set of rules specifies how to handle significant figures in calculations:

1. In addition and subtraction, the answer cannot have more digits to the right of the decimal point than either of the original numbers. Consider these examples:

$$\begin{array}{rl} 89.332 & \\ +\ 1.1 & \longleftarrow \text{one digit after the decimal point} \\ \hline 90.432 & \longleftarrow \text{round off to } 90.4 \end{array}$$

$$\begin{array}{rl} 2.097 & \\ -\ 0.12 & \longleftarrow \text{two digits after the decimal point} \\ \hline 1.977 & \longleftarrow \text{round off to } 1.98 \end{array}$$

2. In multiplication and division, the number of significant figures in the final product or quotient is determined by the original number that has the smallest number of significant figures. The following examples illustrate this rule:

$$2.8 \times 4.5039 = 12.61092 \longleftarrow \text{round off to } 13$$

$$\frac{6.85}{112.04} = 0.0611388789 \longleftarrow \text{round off to } 0.0611$$

3. Keep in mind that exact numbers obtained from definitions or by counting numbers of objects can be considered to have an infinite number of significant figures. For example, the inch is defined to be exactly 2.54 centimeters; that is, 1 in is 2.54 cm

Sample Problem:

Carry out the following arithmetic operations to the correct number of significant figures:

(a) 11,254.1 g 1 0.1983 g
(b) 66.59 L 2 3.113 L
(c) 8.16 m 3 5.1355,
(d) 0.0154 kg 4 88.3 mL
(e) 2.64 3 10^3cm - 3.27 3 10^2 cm

(a)
$$\begin{array}{rl} 11{,}254.1\text{ g} & \\ +\quad 0.1983\text{ g} & \\ \hline 11{,}254.2983\text{ g} & \longleftarrow \text{round off to } 11{,}254.3\text{ g} \end{array}$$

(b)
$$\begin{array}{rl} 66.59\text{ L} & \\ -\ 3.113\text{ L} & \\ \hline 63.477\text{ L} & \longleftarrow \text{round off to } 63.48\text{ L} \end{array}$$

(c) $8.16\text{ m} \times 5.1355 = 41.90568\text{ m} \longleftarrow \text{round off to } 41.9\text{ m}$

(d) $\dfrac{0.0154\text{ kg}}{88.3\text{ mL}} = 0.000174405436\text{ kg/mL} \longleftarrow \text{round off to } 0.000174\text{ kg/mL}$ or 1.74×10^{-4} kg/mL

(e) First we change 3.27×10^2 cm to 0.327×10^3 cm and then carry out the addition $(2.64\text{ cm} + 0.327\text{ cm}) \times 10^3$. Following the procedure in (a), we find the answer is 2.97×10^3 cm.

Lesson 2: Atoms and Molecules

- Explain the difference between a molecular formula and an empirical formula.
- Determine the number of atoms in a molecule from its chemical formula
- Use standard chemical nomenclature to deduce the names of simple inorganic compounds from their formulas or vice versa.

Definition of Terms

Atom The basic unit of an element that can enter into chemical combination.

Atomic Number The number of protons in a particular atom.

Electron A subatomic particle that has a very low mass and carries a single negative electric charge.

Isotopes Atoms of the same element that have different numbers of neutrons.

Mass Number The combined total of protons and neutrons.

Nucleus The central core of an atom.

Neutron A subatomic particle that bears no net electric charge. Its mass is slightly greater than a proton's.

Proton A subatomic particle having a single positive electric charge. The mass of a proton is about 1840 times that of an electron.

The Atomic Theory

- Democritus (Greek philosopher) believed that there was a smallest particle—"atomos" (uncuttable, indivisible)—that made up all of nature.
- Experiments in the eighteenth and nineteenth centuries led to an organized atomic theory by John Dalton in the early 1800s, which explained several laws known at that time:
    1. The law of constant composition
    2. The law of conservation of mass
    3. The law of multiple proportions
- Dalton's Atomic Theory:

1. Elements are composed of extremely small particles called atoms.
2. All atoms of a given element are identical, having the same size, mass, and chemical properties. The atoms of one element are different from the atoms of all other elements.
3. Compounds are composed of atoms of more than one element. In any compound,
the ratio of the numbers of atoms of any two of the elements present is either an
integer or a simple fraction.
4. A chemical reaction involves only the separation, combination, or rearrangement
of atoms; it does not result in their creation or destruction.

Law of Definite Proportion/ Law of Constant Composition

- Different samples of the same compound always contain its constituent elements in the same proportion by mass. Thus, if we were to analyze samples of carbon dioxide gas obtained from different sources, we would find in each sample the same ratio by mass of carbon to oxygen.
- It stands to reason, then, that if the ratio of the masses of different elements in a given compound is fixed, the ratio of the atoms of these elements in the compound also must be constant.
- Also known as the law of definite proportions.
- The elemental composition of a pure substance never varies.
- In a given compound, the relative numbers and kinds of atoms are constant.
- Basis of Dalton's Postulate #4

Law of Multiple Proportions

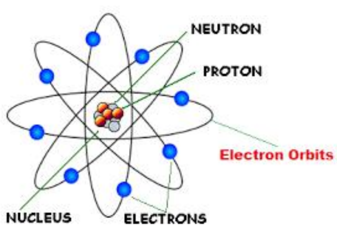
- If two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers. Dalton's theory explains the law of multiple proportions quite simply: Different compounds made up of the same elements differ in the number of atoms of each kind that combine.
- For example, carbon forms two stable compounds with oxygen, namely, carbon monoxide and carbon dioxide. Modern measurement techniques indicate that one atom of carbon combines with one atom of oxygen in carbon monoxide and with two atoms of oxygen in carbon dioxide.mThus, the ratio of oxygen in carbon monoxide to oxygen in carbon dioxide is 1:2. This result is consistent with the law of multiple proportions (Figure 2.2).
- If two elements A and B combine to form more than one compound, the masses of B that can

combine with a given mass of A are in the ratio of small whole numbers.

Law of Conservation of Mass

- Matter can be neither created nor destroyed. Because matter is made of atoms that are unchanged in a chemical reaction, it follows that mass must be conserved as well. Dalton's brilliant insight into the nature of matter was the main stimulus for the rapid progress of chemistry during the nineteenth century.
- The total mass of substances present at the end of a chemical process is the same as the mass of substances present before the process took place.
- Basis of Dalton's Postulate #3
- Can't create matter in a chemical reaction!

The Structure of Atom



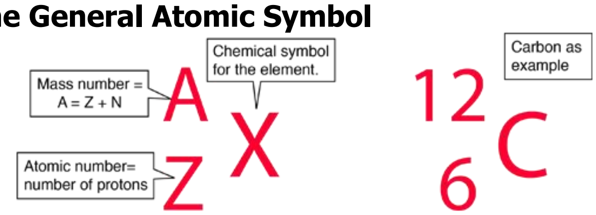
Atom is the basic unit of an element that can enter into chemical combination. Dalton imagined an atom that was both extremely small and indivisible. However, a series of investigations that began in the 1850s and extended into the twentieth century clearly demonstrated that atoms actually possess internal structure; that is, they are made up of even smaller particles, which are called subatomic particles. This research led to the discovery of three such particles—electrons, protons, and neutrons.

The atom is composed of a small, compact core called the nucleus surrounded by a disperse cloud of electrons. The nucleus is composed of two types of particles: protons and neutrons. There is so much space between the electrons and the nucleus that it is impossible to show it to scale in an illustration.

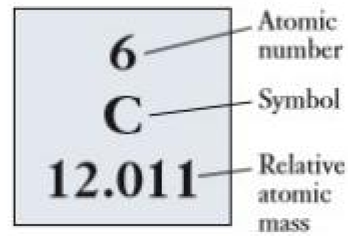
Mass and Charge of Subatomic Particles

Particle	Mass (g)	Charge	
		Coulomb	Charge Unit
Electron*	9.10938×10^{-28}	-1.6022×10^{-19}	-1
Proton	1.67262×10^{-24}	$+1.6022 \times 10^{-19}$	$+1$
Neutron	1.67493×10^{-24}	0	0

The General Atomic Symbol



How to interpret the Periodic Table of Elements



Atomic Mass

This number provides the average mass in amu of an atom of the element. If you look up Carbon in the periodic table inside the back cover of this book, you will find the box shown in Figure 2.2. The atomic mass appears under the symbol: 12.011. But the mass of an atom of carbon-12 is exactly 12 amu, and that of carbon-13 is 13.0036 amu. So the value of 12.011 does not seem to be the mass of any individual atom of carbon.

Naturally occurring isotopes of some selected elements and their atomic masses

Element	Naturally Occurring Isotopes	Atomic Mass (weighted average)
Lithium	$^{6}_{3}Li$ $^{7}_{3}Li$	6.941 amu
Carbon	$^{12}_{6}C$ $^{13}_{6}C$ $^{14}_{6}C$	12.01 amu
Oxygen	$^{16}_{8}O$ $^{17}_{8}O$ $^{18}_{8}O$	16.00 amu
Flourine	$^{19}_{9}F$	19.00 amu
Sulfur	$^{32}_{16}S$ $^{33}_{16}S$ $^{34}_{16}S$ $^{36}_{16}S$	32.07 amu
Potassium	$^{39}_{19}K$ $^{40}_{19}K$ $^{41}_{19}K$	39.10 amu
Copper	$^{63}_{29}Cu$ $^{65}_{29}Cu$	63.55 amu

How are atomic masses defined and determined?

The atomic mass is defined as the average mass of an atom of a particular element.
Carbon has two stable isotopes with masses of 12.0000 and 13.0036 amu, respectively.
So why is the average mass 12.011 and not something closer to 12.5? The answer is that when we take the average mass, we must account for the relative abundance of each isotope.

Sample Problem:

Suppose that we could measure the mass of a 100-atom sample. Based on
the isotopic abundances, we would expect to have 99 atoms of carbon-12 and only a
single atom of carbon-13. In any sample that we can actually weigh, the number of
atoms will be far greater than 100. Even using the best available laboratory balances,
the smallest quantity of matter that can be weighed is about a nanogram, or 10−9 g. A

nanogram of carbon would contain more than 1013 atoms.

Answer:

For such large numbers of atoms, it is safe to assume that the fraction of each isotope present will be determined by the natural abundances. For carbon, the fact that we only need to consider two stable isotopes makes the calculation fairly simple. We can multiply the mass by the fractional abundance to weight each isotope's contribution to the atomic mass.

Carbon-12:	12.0000 × 0.9893	= 11.87
Carbon-13:	13.0036 × 0.0107	= 0.139
Weighted average mass		= 11.87 + 0.139 = **12.01**

The value of **12.011** found in the periodic table is obtained using additional significant figures on the isotopic abundance numbers.

Sample Problem:

The chlorine present in PVC has two stable isotopes. 35Cl with a mass of 34.97 amu makes up 75.77% of the natural chlorine found. The other isotope is 37Cl, whose mass is 36.95 amu. What is the atomic mass of chlorine?

Answer:

First, we calculate the abundance of the chlorine-37 isotope:

Abundance of 37Cl = 100% − 75.77% = 24.23%

35Cl: 34.97 × 0.7577 = 26.50

37Cl: 36.95 × 0.2423 = 8.953

Weighted average mass = 26.50 + 8.953 = 35.45

So the atomic mass of chlorine is 35.45 amu.

Sample Problem:

Magnesium is a macromineral needed in the contraction of muscles and metabolic reactions. Using the data below, calculate the atomic mass for magnesium using the weighted average mass method.

Atomic Symbol	$^{24}_{12}Mg$	$^{25}_{12}Mg$	$^{26}_{12}Mg$
Name	Mg-24	Mg-25	Mg-26
Number of Protons	12	12	12
Number of Electrons	12	12	12
Mass Number	24	25	26
Number of Neutrons	12	13	14
Mass of isotope (amu)	23.99	24.99	25.99
% Abundance	78.70	10.13	11.17

Answer:

$^{24}_{12}Mg$	23.99 × 0.7870	= 18.88
$^{25}_{12}Mg$	24.99 x 0.1013	= 2.531
$^{26}_{12}Mg$	25.99 × 0.1117	= 2.902
Weighted average mass		= 18.88 + 2.531+ 2.902
		= **24.31**

Lesson 3: Periodic Table

- ○ Familiarize the different periods and families or groups in the Periodic Table of Elements.
- ○ Explain the difference between a molecular formula and an empirical formula.
- ○ Determine the number of atoms in a molecule from its chemical formula
- ○ Use standard chemical nomenclature to deduce the names of simple inorganic compounds from their formulas or vice versa.

Definition of Terms

Allotrope is one of two or more distinct forms of an element.

Anion A negatively charged atom or group of atoms.

Cation A positively charged atom or group of atoms.

Group/Family The elements in a vertical column of the periodic table.

Ion An atom or a group of atoms that has a net positive or negative charge.

Metalloid An element with properties intermediate between those of metals and nonmetals.

Metals Elements that are good conductors of heat and electricity and have the tendency to form positive ions in ionic compounds.

Nonmetals Elements that are usually poor conductors of heat and electricity.

Period A horizontal row of the periodic table.

Periodic Law: when properly arranged, the elements display a regular and periodic variation in their chemical properties.

Periodic Table A chart in which elements having similar chemical and physical properties are grouped together.

History and development of the periodic table
Johann Dobereiner (1780 - 1849)

- In 1829, a German chemist, Johann Dobereiner (1780 - 1849), placed various groups of three elements into groups called triads. One such triad was lithium, sodium, and potassium. Triads were based on both physical as well as chemical properties.
- Dobereiner found that the atomic masses of these three elements, as well as other triads, formed a pattern.
- While Dobereiner's system would pave the way for future ideas, a limitation of the triad system was that not all of the known elements could be classified in this way.

John Newlands (1838 - 1898)

- John Newlands (1838 - 1898) ordered the elements in increasing order of atomic mass and noticed that every eighth element exhibited similar properties. He called this relationship the "Law of Octaves". Unfortunately, there were some elements that were missing and the law did not seem to hold for elements that were heavier than calcium.
- Newlands' work was largely ignored and even ridiculed by the scientific community in his day. It was not until years later that another, more extensive periodic table effort would gain much greater acceptance and the pioneering work of John Newlands would be appreciated.

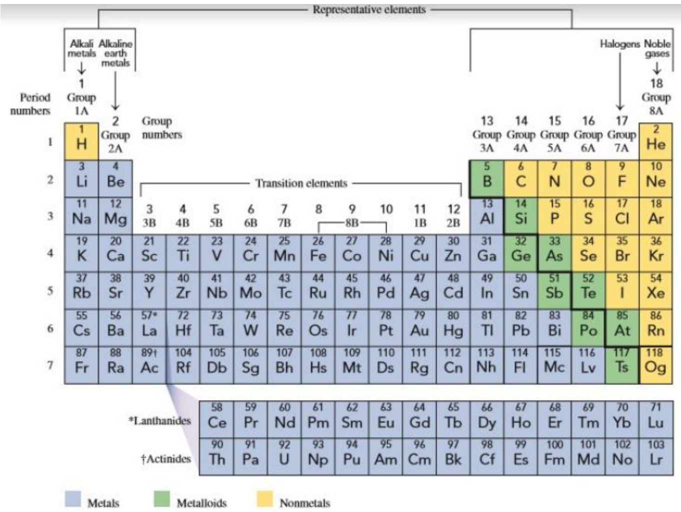
Dmitri Mendeleev (1836 - 1907)

- In 1869, Russian chemist and teacher Dmitri Mendeleev (1836 - 1907) published a periodic table of the elements.
- The following year, German chemist Lothar Meyer independently published a very similar table.
- Mendeleev is generally given more credit than Meyer because his table was published first and because of several key insights that he made regarding the table.

The Modern Periodic Table of Elements

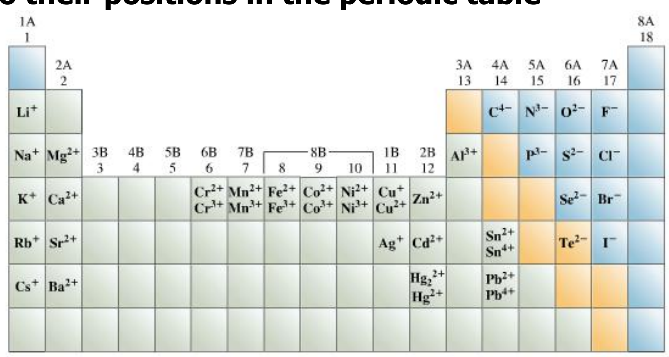
Column 1. Alkali Metals (Group 1A)
Column 2. Alkali Earths (Group 2A)
Columns 3- 12. Transition Metals

Column 17. Halogens (Group 7A)
Column 18. Noble Gases (Group 8A)
Row 8. Lanthanides
Row 9. Actinides



The elements can be divided into three categories—metals, nonmetals, and metalloids. A metal is a good conductor of heat and electricity while a nonmetal is usually a poor conductor of heat and electricity. A metalloid has properties that are intermediate between those of metals and nonmetals.

Common monatomic ions arranged according to their positions in the periodic table



Chemical Formula

<u>Molecular Formulas</u>

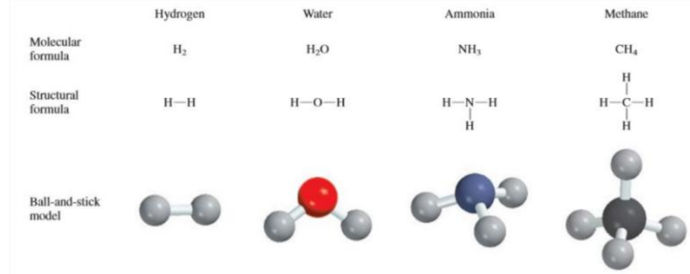
shows the exact number of atoms of each element in the smallest unit of a substance.

<u>Empirical Formulas</u>

tells us which elements are present and the simplest whole-number ratio of their atoms, but not necessarily the actual number of atoms in a given molecule.

Empirical formulas are the simplest chemical formulas; they are written by reducing the subscripts in the molecular formulas to the smallest possible whole numbers. Molecular formulas are the true formulas of molecules.

Molecular and structural formula of four common molecules



Sample Problem:

Write the empirical formulas for the following molecules: (a) acetylene (C2H2), which is used in welding torches; (b) glucose (C6H12O6), a substance known as blood sugar; and (c) nitrous oxide (N2O), a gas that is used as an anesthetic gas ("laughing gas") and as an aerosol propellant for whipped creams.

Solution:

To write the empirical formula, which is the simplest form, the subscripts in the molecular formula must be converted to the smallest possible whole number.

Molecule	Molecular Formula	Divide the subscript by	Empirical Formula
a. acetylene	C_2H_2	2	**CH**
b. glucose	$C_6H_{12}O_6$	6	**CH_2O**
c. nitrous oxide	N_2O	1	**N_2O**

Naming Compounds

- Organic compounds contain carbon, usually in combination with elements such as hydrogen, oxygen, nitrogen, and sulfur (H, O, N, S). All other compounds are classified as inorganic compounds
- To organize and simplify our venture into naming compounds, we can divide inorganic compounds into four categories: ionic compounds, molecular compounds, acids and bases, and hydrates.

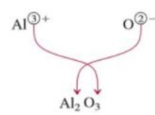
Naming Compounds: For Ionic Compounds

- For ionic compounds to be electrically neutral, the sum of the charges on the cation and anion in each formula unit must be zero. If the charges on the cation and anion are numerically different, we apply the following rule to make the formula electrically neutral: The subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation. If the charges are numerically equal, then no subscripts are necessary. *This rule follows from the fact that because the formulas of ionic compounds are usually empirical formulas, the subscripts must always be reduced to the smallest ratios.*
- Let us consider some examples

1. Potassium Bromide. The potassium cation K+ and the bromine anion Br- combine to form the ionic compound potassium bromide. The sum of the charges is +1 + (-1) = 0, so no subscripts are necessary. The formula is KBr.

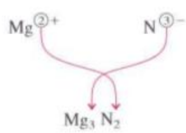
2. Zinc Iodide. The zinc cation Zn2+ and the iodine anion I- combine to form zinc iodide. The sum of the charges of one Zn2+ ion and one I- ion is +2 + (-1) = +1. To make the charges add up to zero we multiply the -1 charge of the anion by 2 and add the subscript "2" to the symbol for iodine. Therefore the formula for zinc iodide is ZnI2.

3. Aluminum Oxide. The cation is Al3+ and the oxygen anion is O2-. The following diagram helps us determine the subscripts for the compound formed by the cation and the anion. The sum of the charges is 2(+3) + 3(-2) = 0. Thus the formula for aluminum oxide is Al2O3.



Example 3.2 Write the formula of magnesium nitride, containing the Mg^{2+} and N^{3-} ions.

SOLUTION *(See Example 2.4 on Text)*



Names and Formulas of Some Common Inorganic Cations and Anions

Cation	Anion
aluminum (Al^{3+})	bromide (Br^-)
ammonium (NH_4^+)	carbonate (CO_3^{2-})
barium (Ba^{2+})	chlorate (ClO_3^-)
cadmium (Cd^{2+})	chloride (Cl^-)
calcium (Ca^{2+})	chromate (CrO_4^{2-})
cesium (Cs^+)	cyanide (CN^-)
chromium(III) or chromic (Cr^{3+})	dichromate ($Cr_2O_7^{2-}$)
cobalt(II) or cobaltous (Co^{2+})	dihydrogen phosphate ($H_2PO_4^-$)
copper(I) or cuprous (Cu^+)	fluoride (F^-)
copper(II) or cupric (Cu^{2+})	hydride (H^-)
hydrogen (H^+)	hydrogen carbonate or bicarbonate (HCO_3^-)
iron(II) or ferrous (Fe^{2+})	hydrogen phosphate (HPO_4^{2-})
iron(III) or ferric (Fe^{3+})	hydrogen sulfate or bisulfate (HSO_4^-)
lead(II) or plumbous (Pb^{2+})	hydroxide (OH^-)
lithium (Li^+)	iodide (I^-)
magnesium (Mg^{2+})	nitrate (NO_3^-)
manganese(II) or manganous (Mn^{2+})	nitride (N^{3-})
mercury(I) or mercurous (Hg_2^{2+})*	nitrite (NO_2^-)
mercury(II) or mercuric (Hg^{2+})	oxide (O^{2-})
potassium (K^+)	permanganate (MnO_4^-)
rubidium (Rb^+)	peroxide (O_2^{2-})
silver (Ag^+)	phosphate (PO_4^{3-})
sodium (Na^+)	sulfate (SO_4^{2-})
strontium (Sr^{2+})	sulfide (S^{2-})
tin(II) or stannous (Sn^{2+})	sulfite (SO_3^{2-})
zinc (Zn^{2+})	thiocyanate (SCN^-)

Molecular Compounds

We place the name of the first element in the formula first, and the second element is named by adding -ide to the root of the element name.

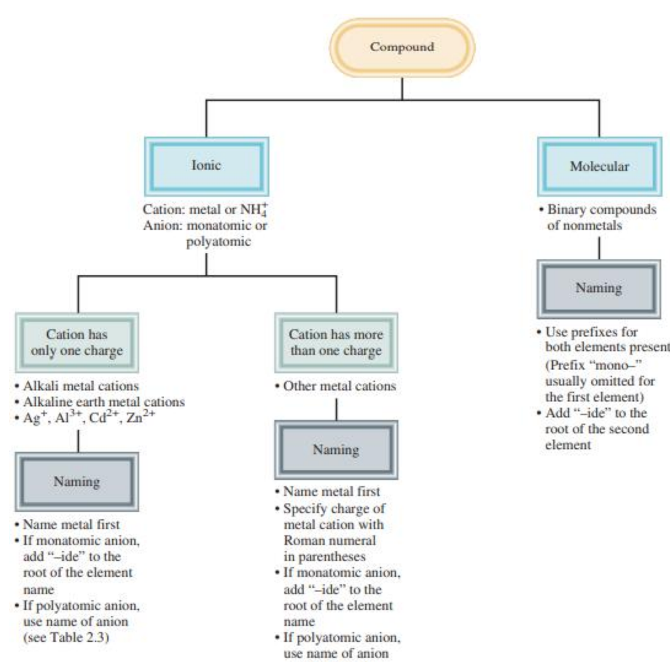
Greek Prefixes Used in Naming Molecular Compounds

Prefix	Meaning
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

HCl hydrogen chloride
HBr hydrogen bromide
SiC silicon carbide
CO carbon monoxide
CO2 carbon dioxide
SO2 sulfur dioxide
SO3 sulfur trioxide
NO2 nitrogen dioxide
N2O4 dinitrogen tetroxide

The following guidelines are helpful in naming compounds with prefi xes:

- The prefix "mono-" may be omitted for the first element. For example, PCl3 is named phosphorus trichloride, not monophosphorus trichloride. Thus, the absence of a prefix for the first element usually means there is only one atom of that element present in the molecule.
- For oxides, the ending "a" in the prefi x is sometimes omitted. For example, N2O4 may be called dinitrogen tetroxide rather than dinitrogen tetraoxide.



Module 4A: Chemical Equations and Stoichiometry

- ◯ Explain balancing a chemical equation as an application of the law of conservation of mass.
- ◯ Write balanced chemical equations for simple reactions, given either an unbalanced equation or a verbal description.
- ◯ Interpret chemical equations in terms of both moles and molecules.
- ◯ Interconvert between mass, number of molecules, and number of moles.

Definition of Terms:

Chemical Equation uses chemical symbols to show what happens during a chemical reaction.

Chemical Reaction A process in which a substance (or substances) is changed into one or more new substances.

Empirical Formula The smallest whole-number ratio of atoms present in a compound; also called empirical formula.

Excess Reactant The reactants present in quantities greater than necessary to react with the quantity of the limiting reagent.

Formula Weight

- The mass, in atomic mass units, of one formula unit of a substance.
- Numerically equal to the mass, in grams, of one mole of the substance (see Molar mass).
- This number is obtained by adding the atomic weights of the atoms specified in the formula.

Limiting Reactant The reactant that is completely consumed in a reaction. The available amount of it determines the maximum possible reaction yield.

Molar mass The mass of substance in one mole of the substance; numerically equal to the formula weight of the substance.

Mole

- 6.022 x 1023 (Avogadro's number of) formula units (or molecules, for a molecular substance) of the substance under discussion.
- The mass of one mole, in grams, is numerically equal to the formula (molecular) weight of the substance.

Molecular Weight

- The mass, in atomic mass units, of one molecule of a nonionic (molecular) substance. Numerically equal to the mass, in grams, of one mole of such a substance.
- This number is obtained by adding the atomic weights of the atoms specified in the formula.

Molecule The smallest particle of an element or compound that can have a stable independent existence. Percent Composition The mass percentage of each element in a compound.

Product is the substance formed as a result of a chemical reaction.

Reactants are the starting materials in a chemical reaction.

Stoichiometric Amount Amount that is in the proportions indicated by the balanced equation.

Atomic Mass

- Atomic Mass (sometimes called atomic weight) is the mass of the atom in atomic mass units (amu).
- One atomic mass unit is defined as a mass exactly equal to one-twelfth the mass of one carbon-12 atom.
- Carbon-12 is the carbon isotope that has six protons and six neutrons. Setting the atomic mass of carbon-12 at 12 amu provides the standard for measuring the atomic mass of the other elements.
- Atomic masses do not convert easily to grams
- They can't be weighed (they are too small)

Why do we need to use atomic mass?

✓ Masses give information about # of p+, n0, e−

✓ It is useful to know relative mass

✓ It is useful to associate atomic mass with a mass in grams. It has been found that 1 g H, 12 g C, or 23 g Na have 6.02 x 1023 atoms

THE MOLE (n)

Avogadro's Number and the Molar Mass of an Element

- In chemistry, particles such as atoms, molecules, and ions are counted by the mole (abbreviated mol in calculations), which contains 6.022 x 1023 items. This value, known as Avogadro's number, is a very big number because atoms are so small that it takes an extremely large number of atoms to provide a sufficient amount to weigh and use in chemical reactions.

- Avogadro's number is named for Amedeo Avogadro (1776–1856), an Italian physicist.
- 6.02 x 1023 is a "mole" or "Avogadro's number" "mol" is used in equations, "mole" is used in writing; one gram = 1 g, one mole = 1 mol.

6.022 x 1023 = 602 200 000 000 000 000 000 000

Using Avogadro's Number as a Conversion Factor

$$\frac{6.022 \times 10^{23}\,\text{particles}}{1\ mol} \quad \text{and} \quad \frac{1\ mol}{6.022 \times 10^{23}\,\text{particles}}$$

Example:

How many molecules are present in 1.75 mol of carbon dioxide?

$$1.75\ mol\,\cancel{CO_2}\ x\ \frac{6.022 \times 10^{23}\,\text{molecules}\ CO_2}{1\ mol\,\cancel{CO_2}} = \mathbf{1.05\ x\ 10^{24} molecules\ CO_2}$$

Moles of Elements in a Chemical Compound

$C_9H_8O_4$

	Carbon	Hydrogen	Oxygen
Atoms in 1 molecule	9 atoms of C	8 atoms of H	4 atoms of O
Moles of each element in 1 mol	9 mol of C	8 mol of H	4 mol of O

Molar Mass (MM)

To determine the molar mass of a compound, multiply the molar mass of each element by its subscript in the formula and add the results.

- Molar mass is the quantity in grams that equals the atomic mass of that element. We are counting 6.022 x 1023 atoms of an element when we weigh out the number of grams equal to its molar mass.
- For example, carbon has an atomic mass of 12.01 on the periodic table. This means 1 mol of carbon atoms has a mass of 12.01 g. Then to obtain 1 mol of carbon atoms, we would need to weigh out 12.01 g of carbon. Thus, the molar mass of carbon is found by looking at its atomic mass on the periodic table.

Example

Calculate the molar mass for lithium carbonate, Li2CO3, used to treat bipolar disorder.

$$2\ mol\ Li\left(\frac{6.941\ g\ Li}{1\ mol\ Li}\right) = 13.88\ g\ Li$$

$$1\ mol\ C\left(\frac{12.01\ g\ C}{1\ mol\ C}\right) = 12.01\ g\ C$$

$$3\ mol\ O\left(\frac{16.00\ g\ O}{1\ mol\ O}\right) = 48.00\ g\ O$$

Molar mass of Li2CO3 is the sum of the masses of each element:

$$13.88\ g\ Li + 12.01\ g\ C + 48.00\ g\ O = \mathbf{73.89\ g}$$

- The molar mass of an element is one of the most useful conversion factors in chemistry because it converts moles of a substance to grams, or grams to moles. For example, 1 mol of silver has a mass of 107.9 g. To express molar mass of Ag as an equality, we write

1 mol Ag = 107.9 g Ag

From this equality for the molar mass, two conversion factors can be written as:

$$\frac{107.9\ g\ Ag}{1\ mol\ Ag} \quad \text{and} \quad \frac{1\ mol\ Ag}{107.9\ g\ Ag}$$

Conversion between moles and grams:

Example

Silver metal is used in the manufacture of tableware, mirrors, jewelry, and dental alloys. If the design for a piece of jewelry requires 0.750 mol of silver, how many grams of silver are needed?

Solution:

Conversion factor:

$$\frac{107.9\ g\ Ag}{1\ mol\ Ag} \quad \text{and} \quad \frac{1\ mol\ Ag}{107.9\ g\ Ag}$$

Just select the proper conversion factor to convert moles to grams.

$$0.750\ \cancel{mol\ Ag}\ \ x\ \ \frac{107.9\ g\ Ag}{1\ \cancel{mol\ Ag}} = \mathbf{80.9\ g\ Ag}$$

Conversion between mass and moles of a compound:

Example

A salt shaker contains 73.7 g of NaCl. How many moles of NaCl are present?

Solution:

First, determine the Molar Mass of NaCl and form conversion factors.

MM is 58.44 g/mol

$$\frac{58.44\ g\ NaCl}{1\ mol\ NaCl} \quad and \quad \frac{1\ mol\ NaCl}{58.44\ g\ NaCl}$$

From here, we now can convert grams to moles.

$$73.7\ \cancel{g\ NaCl}\ \ x\ \ \frac{1\ mol\ NaCl}{58.44\ \cancel{g\ NaCl}} = \mathbf{1.26\ mol\ NaCl}$$

Converting between grams of compound and grams of element:

Example:

Hot packs are used to reduce muscle aches, inflammation, and muscle spasms. A hot pack consists of a bag of water and an inner bag containing 10.2 g of CaCl2. When the bag is broken, the CaCl2 dissolves in the water and heat is released. How many grams of Cl are in the CaCl2 in the inner bag?
Solution:
conversion factors for molar mass and mole factors are obtained for this example:

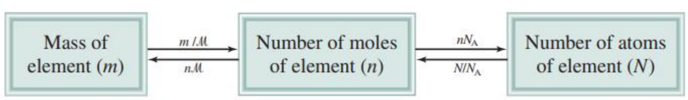
$$1\ mol\ CaCl_2 = 110.98\ g\ of\ CaCl_2 \qquad \frac{110.98\ g\ CaCl_2}{1\ mol\ CaCl_2}\ \ and\ \ \frac{1\ mol\ CaCl_2}{110.98\ g\ CaCl_2}$$

$$1\ mol\ CaCl_2 = 2\ mol\ Cl \qquad \frac{2\ mol\ Cl}{1\ mol\ CaCl_2}\ \ and\ \ \frac{1\ mol\ CaCl_2}{2\ mol\ Cl}$$

$$1\ mol\ Cl = 35.45\ g\ Cl \qquad \frac{35.45\ g\ Cl}{1\ mol\ Cl}\ \ and\ \ \frac{1\ mol\ Cl}{35.45\ g\ Cl}$$

We can proceed to the conversion:

$$10.2\ g\ \cancel{CaCl_2}\ x\ \frac{1\ \cancel{mol\ CaCl_2}}{110.98\ g\ \cancel{CaCl_2}}\ x\ \frac{2\ \cancel{mol\ Cl}}{1\ \cancel{mol\ CaCl_2}}\ x\ \frac{35.45\ g\ Cl}{1\ \cancel{mol\ Cl}} = \mathbf{6.52\ g\ of\ Cl}$$



Mass Percent Composition

we can calculate its mass percent composition or mass percent, which is the mass of an element divided by the total mass of the compound and multiplied by 100%.

$$Mass\ Percent\ of\ an\ Element\ = \frac{mass\ of\ an\ element}{total\ mass\ of\ the\ compound}\ x\ 100\%$$

$$Mass\ Percent\ composition\ = \frac{mass\ of\ each\ element}{molar\ mass\ of\ the\ compound}\ x\ 100\%$$

Example

The odor of pears is due to the compound propyl acetate, which has a formula of C5H10O2. What is the mass percent composition of propyl acetate?
Solution
Determine the total mass of each element in the molar mass of a formula.

$$5\ mol\ C\ x\ \frac{12.01\ gC}{1mol\ C} = 60.05\ g\ of\ C$$

$$10\ mol\ H\ x \frac{1.008\ \ g\ H}{1\ mol\ H} = 10.08\ g\ of\ \ H$$

$$2\ mol\ O\ x \frac{16.00\ g}{1\ mol\ O} = 32.00\ g\ \ O$$

Use the formula above to find the mass % of each element:

$$Mass\ \%\ C = \frac{60.05\ g\ of\ C}{102.13\ g\ C_5H_{10}O_2}\ x\ 100 = \mathbf{58.80\%\ C}$$

$$Mass\ \%\ H = \frac{10.08\ g\ of\ \ H}{102.13\ g\ C_5H_{10}O_2}\ x\ 100 = \mathbf{9.870\%\ H}$$

$$Mass\ \%\ O\ = \frac{32.00\ g\ \ O}{102.13\ g\ C_5H_{10}O_2}\ x\ 100\ = \mathbf{31.33\%\ O}$$

The total mass percent for all the elements in the compound should equal 100%. In some cases, because of rounding off, the sum of the mass percents may not total exactly 100%.

Example :

Ascorbic acid (vitamin C) cures scurvy. It is composed of 40.92 percent carbon (C), 4.58 percent hydrogen (H), and 54.50 percent oxygen (O) by mass. Determine its empirical formula.
Solution:
Basis: 100 g of Ascorbic Acid

$$\boldsymbol{n_C} = 40.92\%\ 100g\ x \frac{1\ mol\ C}{12.01\ g\ C} = 3.407\ mol\ C$$

$$\boldsymbol{n_H} = 4.58\%\ 100g\ \ x \frac{1\ mol\ H}{1.008\ g\ H} = 4.540\ mol\ H$$

$$\boldsymbol{n_O} = 54.50\%\ 100g\ \ x \frac{1\ mol\ O}{16.00\ g\ O} = 3.406\ mol\ H$$

To simplify the subscript is to divide all the subscripts by the smallest subscript. In this case, 3.406

$$\boldsymbol{n_C} = 3.407\ mol\ C\ \div\ 3.406\ mol\ \ = 1$$

$$\boldsymbol{n_H} = 4.540\ mol\ H\ \div\ 3.406\ mol\ \ 1.33$$

$$\boldsymbol{n_O} = 3.406\ mol\ H\ \div\ 3.406\ mol = 1$$

Next, we need to convert 1.33, the subscript for H, into an integer. This can be done by a trial-and-error procedure. From which, you can see that only by multiplying each subscript by 3 can we acquire an integer or whole number. Therefore the empirical formula for ascorbic acid is: C3H4O3

Calculating the Molecular Formula

Example

Melamine, which is used to make plastic items such as dishes and toys, contains 28.57% C, 4.80% H, and 66.64% N. If the experimental molar mass is 125 g, what is the molecular formula of melamine?

Solution:

1. Obtain the empirical formula and calculate the empirical formula mass

Basis: 100 g of this compound, therefore, there are 28.57 g of C, 4.80 g of H, and 66.64 g of N.

$$\boldsymbol{n_C} = 28.57\ g\ \ x\frac{1\ mol\ C}{12.01\ g\ C} = 2.38\ mol\ C$$

$$\boldsymbol{n_H} = 4.80\ g\ \ x\frac{1\ mol\ H}{1.008\ g\ H} = 4.76\ mol\ H$$

$$\boldsymbol{n_N} = 66.64\ g\ \ x\frac{1\ mol\ N}{14.01\ g\ N} = 4.76\ mol\ N$$

2. Divide the moles of each element by the smallest number of moles, 2.38, to obtain the subscripts of each element in the formula.

$$\frac{2.38\ mol\ C}{2.38} = 1\ mol\ C \qquad \frac{4.76\ mol\ H}{2.38} = 2\ mol\ H \qquad \frac{4.76\ mol\ N}{2.38} = 2mol\ N$$

the empirical formula for melamine as CH2N2.

3. Now we calculate the molar mass for this empirical formula as follows:

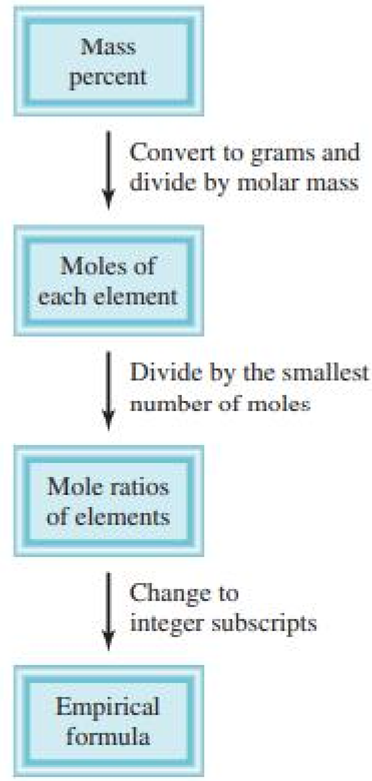
Empirical Formula Mass = (1 x 12.01) + (2 x 1.008) + (2 x 14.01) = 42.05 g

4. Divide the molar mass by the empirical formula mass

$$\frac{molar\ mass\ of\ melamine}{empirical\ formula\ mass\ of\ \text{CH2N2}} = \frac{125\ g}{42.05\ g} = 2.97 \approx 3$$

5. Multiply the empirical formula by this factor to obtain the molecular formula. Because the experimental molar mass is close to 3 times the empirical formula mass, the subscripts in the empirical formula are multiplied by 3 to give the molecular formula, **C3H6N6**.

Procedure for calculating the empirical formula of a compound from its percent compositions



Lesson 4B: Chemical Equations and Stoichiometry

CHEMICAL EQUATIONS

A chemical change occurs when a substance is converted into one or more new substances. For example, when silver tarnishes, the shiny silver metal (Ag) reacts with sulfur (S) to become the dull, black substance we call tarnish (Ag2S).

WRITING A CHEMICAL EQUATION

Consider what happens when hydrogen gas (H 2) burns in air (which contains oxygen, O 2) to form water (H 2 O). This reaction can be represented by the chemical equation

$$H_2 + O_2 \longrightarrow H_2O$$

where the "plus" sign means "reacts with" and the arrow means "to yield." Thus, this symbolic expression can be read: "Molecular hydrogen reacts with molecular oxygen to yield water." The reaction is assumed to proceed from left to right as the arrow indicates Generally, each formula in an equation is followed by an abbreviation, in parentheses, that gives the physical state of the substance: solid (s), liquid (l), or gas (g). If a substance is dissolved in water, it is in an aqueous (aq) solution. The delta sign (Δ) indicates that heat was used to start the reaction.

DATE ____________

Evidence of Chemical Reactions

- Visible evidence of chemical reactions include
  - i) color change
  - ii) the formation of a solid (precipitate) in a previously clear solution
  - iii) the formation of a gas when you add a substance to a solution
  - iv) a change in energy (heat, light)
- Chemical reactions may occur w/o any obvious signs, yet chemical analysis may show that a reaction has indeed occured

Chemical Equations

- Chemical reactions are represented by chemical equations
- We often specify the state of each reactant or product in parentheses next to the formula.

abbreviation	state
(g)	gas
(l)	liquid
(s)	solid
(aq)	aqueous

↓

indicates that a substance is dissolved in water. when a substance dissolves in water, the mixture is called solution

Premiere Notes

DATE

Chemical Equations

- The reaction occuring in a natural gas flame is methane reacting with oxygen to form carbon dioxide and water

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Balancing Equations

Law of Conservation of Mass

- In a chemical (equation) reaction, matter is neither created nor destroy
- If an equation obeys the Law of Conservation, it is balanced

Types of chemical reaction

1. Synthesis (the get together)

$$A + B \rightarrow AB$$

2. Decomposition (the break up)

$$AB \longrightarrow A + B$$

2. Single Replacement (the cheater)

$$AB + C \rightarrow AC + B$$

4. Double Replacement (the swap)

$$AB + CD \rightarrow AD + BC$$

5. Combustion (everyone loves O_2)

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DATE

Stoichiometry

- Stoichiometry is the study of substances and their relationship to each other in chemical reactions
- A balanced chemical equation provides several important information about the reactants and products in a chemical reaction

MOLAR RATIOS

$$1N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

1 molecule	3 molecules	2 molecules
100 molecules	300 molecules	200 molecules
1×10^6 molecules	3×10^6 molecules	2×10^6 molecules
1 mole	3 mole	2 mole

Example: Determine each mole ratio below based on the reaction shown

$$2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$$

$$\frac{\text{mole } O_2}{\text{mole } CO_2} = \frac{18}{3}$$

$$\frac{\text{mol } C_4H_{10}}{\text{mol } H_2O} = \frac{2}{10}$$

Stoichiometry

Stoichiometry allows us to predict products that form in a reaction based on amount of reactant.

MOLE TO MOLE CALCULATIONS

relates moles of reactants and products in a balanced chemical equation

moles of compound A $\xleftrightarrow{\text{molar ratio}}$ moles of compound B

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DATE

mole to mole calculations

How many moles of ammonia can be produced from 32 moles of hydrogen? (Assume excess N_2 present)

$$1N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

$$\frac{32 \text{ moles } \cancel{H_2}}{} \left| \frac{2 \text{ moles } NH_3}{3 \text{ moles } \cancel{H_2}} \right. = 21.33 \text{ moles } NH_3$$

In one experiment, 6.80 mol of ammonia are prepared. How many moles of hydrogen were used up in this experiment?

$$\frac{6.80 \cancel{\text{ moles } NH_3}}{} \left| \frac{3 \text{ mol } H_2}{2 \cancel{\text{mol } NH_3}} \right. = 10.2 \text{ mol } H_2$$

$$1N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

MASS - MOLE CALCULATION

relates moles and mass of reactants or products in a balanced chemical equation

mass of compound A

$\updownarrow$ molar mass

moles of compound A $\xleftrightarrow{\text{molar ratio}}$ **moles of compound B**

mass to mole calculation

How many moles of ammonia can be produced from the reaction of 125 g of nitrogen?

$$1N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

$$\frac{125 \text{ g } \cancel{N_2}}{} \left| \frac{1 \cancel{\text{mole } N_2}}{28 \cancel{\text{ g } N_2}} \right| \frac{2 \text{ mol } NH_3}{1 \cancel{\text{mole } N_2}} = 8.93 \text{ mol } NH_3$$

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DATE ______

MASS - MASS CALCULATIONS

relates mass of reactants and products in a balanced chemical equation

$$\begin{array}{ccc} \textbf{mass of compound A} & & \textbf{mass of compound B} \\ \updownarrow \text{molar mass} & & \text{molar mass} \updownarrow \\ \textbf{moles of compound A} & \xleftrightarrow[\text{molar ratio}]{} & \textbf{moles of compound B} \end{array}$$

mass to mass calculation

What mass of Carbon Dioxide will be produced from the reaction of 175 g of propane, as shown

$$1\,C_3H_8\,(g) + 5O_2\,(g) \longrightarrow 3CO_2\,(g) + 4H_2O\,(g)$$

process:

mass of C_3H_8 (g) → moles of C_3H_8 → moles CO_2 → mass CO_2

175 g C_3H_8	1 mole C_3H_8	3 mole CO_2	44 g CO_2	= 525 g CO_2
	44 g C_3H_8	1 mole C_3H_8	1 mol CO_2	

How many grams of O_2 are required to react with 22.5 g of C_7H_{16}?

$$C_7H_{16}\,(l) + 11O_2\,(g) \longrightarrow 7CO_2\,(g) + 8H_2O\,(g)$$

22.5 g C_7H_{16}	1 mole C_7H_{16}	11 mole O_2	32 g O_2	= 79.2 g O_2
	100 g C_7H_{16}	1 mole C_7H_{16}	1 mol O_2	

Limiting Reactant

- When 2 or more reactants are combined in non-stoichiometric ratios, the amount of product produced is limited by the reactant that is not in excess
- This reactant is referred to as limiting reagent/reactant
- when doing stoichiomentric problems of this type, the limiting reactant must be determined first before proceeding w/ calculations

DATE ____________

$A + B \rightarrow C$

A is LR $\longrightarrow$ calculate amount of C] lower value is
B is LR $\longrightarrow$ calculate amount of C] the correct

limiting reactant

A fuel mixture used in the early days of rocketry was a mixture of N_2H_4 and N_2O_4, as shown below. How many grams of N_2 gas is produced when 100 g of N_2H_4 and 200 g of N_2O_4 are mixed?

$$2N_2H_4(l) + 1N_2O_4(l) \rightarrow 3N_2(g) + 4H_2O(g)$$

100g N_2H_4	1 mol N_2H_4	3 mol N_2	28 g N_2	= 131.25 g N_2
	32 g N_2H_4	2 mol N_2H_4	1 mol N_2	

200 g N_2O_4	1 mol N_2O_4	3 mol N_2	28 g N_2	= 91.30 g N_2
	92 g N_2O_4	2 mol N_2O_4	1 mol N_2	↓ limiting reactant

How many grams of AgBr can be produced when 50.0 g of $MgBr_2$ is mixed with 100.0 g of $AgNO_3$, as shown

$$MgBr_2 + 2AgNO_3 \rightarrow 2AgBr + Mg(NO_3)_2$$

50.0 g $MgBr_2$	1 mol $MgBr_2$	2 mol AgBr	187 g AgBr	= 101.63 g AgBr ↑ limiting reactant
	184 g $MgBr_2$	1 mol $MgBr_2$	1 mol AgBr	

100.0 g $AgNO_3$	1 mol $AgNO_3$	2 mol AgBr	187 g AgBr	= 110 g AgBr
	170 g $AgNO_3$	2 mol $AgNO_3$	1 mol AgBr	

DATE

Percent Yield

- The amount of product calculated through stoichiometric ratios are the maximum amount product can be produced during the reaction, and is thus called theoretical yield
- The actual yield of a product in a chemical reaction is the actual amount obtained from the reaction

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

percent yield

In an experiment forming ethanol, the theoretical yield is 50.0 g and the actual yield is 46.8 g. What is the percent yield for this reaction

$$\text{Percent yield} = \frac{46.8\,g}{50.0\,g} \times 100$$

$$= 93.6\%$$

Silicon carbide can be formed from the reaction of sand (SiO_2) with carbon as shown below:

$$1SiO_2\,(s) + 3C\,(s) \rightarrow 1SiC\,(s) + 2CO\,(g)$$

When 100g of sand are processed, 51.4g of SiC is produced. What is the percent yield of SiC in this run?

100g SiO_2	1 mol SiO_2	1 mol SiC	40g SiC	= 66.67g
	60 g SiO_2	1mol SiO_2	1mol SiC	

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

$$= \frac{51.4\,g}{66.67\,g} \times 100$$

$$= 77.10\%$$

Premiere Notes

Lesson 5: Energy

- ◯ Identify energy as potential or kinetic; convert between units of energy.
- ◯ Calculate the specific heat for a substance.
- ◯ Use specific heat to calculate heat loss or gain.
- ◯ Define state functions and explain their importance.
- ◯ Explain the economic importance of conversions between different forms of energy and the inevitability of losses in this process

ENERGY

- "Energy" is a much-used term that represents a rather abstract concept.
- energy is known and recognized by its effects. It cannot be seen, touched, smelled, or weighed.
- All forms of energy are capable of doing work (that is, of exerting a force over a distance), but not all of them are equally relevant to chemistry.
- is usually defined as the capacity to do work and work as "force x distance," butthere are other kinds of work.
- Forms of energy:

1. Kinetic energy
   - ➢ the energy produced by a moving object—is one form of energy that is of particular interest to chemists. Others include radiant energy, thermal energy, chemical energy, and potential energy.
2. Potential energy
   - ➢ is energy available by virtue of an object's position.
   - ➢ For instance,because of its altitude, a rock at the top of a cliff has more potential energy and will make a bigger splash if it falls into the water below than a similar rock located partway down the cliff.
   - ➢ Chemical energy can be considered a form of potential energy because it is associated with the relative positions and arrangements of atoms within a given substance.
3. Chemical energy
   - ➢ is energy stored in the bonds of atoms and molecules.
   - ➢ Batteries, biomass, petroleum, natural gas, and coal are examples of chemical energy.
   - ➢ Chemical energy is converted to thermal energy when people burn wood in a fireplace or burn gasoline in a car's engine.
4. Mechanical energy
   - ➢ is energy stored in objects by tension.
   - ➢ Compressed springs and stretched rubber bands are examples of stored mechanical energy.
5. Nuclear energy
   - ➢ is energy stored in the nucleus of an atom—the energy that holds then nucleus together. Large amounts of energy can be released when the nuclei are combined or split apart.
6. Gravitational energy
   - ➢ is energy stored in an object's height. The higher and heavier the object, the more gravitational energy is stored.
   - ➢ When a person rides a bicycle down a steep hill and picks up speed, the gravitational energy is converting to motion energy.
   - ➢ Hydropower is another example of gravitational energy, where gravity forces water down through a hydroelectric turbine to produce electricity.
7. Radiant energy
   - ➢ is electromagnetic energy that travels in transverse waves.
   - ➢ Radiant energy includes visible light, x-rays, gamma rays, and radio waves.
   - ➢ Light is one type of radiant energy. Sunshine is radiant energy, which provides the fuel and warmth that make life on earth possible.
8. Thermal energy or heat
   - ➢ is the energy that comes from the movement of atoms and molecules in a substance.
   - ➢ Heat increases when these particles move faster. Geothermal energy is the thermal energy in the earth.
9. Sound
   - ➢ is the movement of energy through substances in longitudinal (compression/rarefaction) waves.
   - ➢ Sound is produced when a force causes an object or substance to vibrate. The energy is transferred through the substance in a wave. Typically, the energy in sound is smaller than in other forms of energy.
10. Electrical energy
    - ➢ is delivered by tiny charged particles called electrons, typically moving through a wire.
    - ➢ Lightning is an example of electrical energy in nature.

All forms of energy can be converted (at least in principle) from one form to another. Although energy can assume many different forms that are interconvertible, scientist have concluded that energy can be neither destroyed nor created. When one form of energy disappears, some other form of energy (of equal magnitude) must appear, and vice versa. This principle is summarized by the Law of Conservation of Energy: the total quantity of energy in the universe is assumed constant.

Sources of Energy

1. Non-renewable

- Most of our energy is nonrenewable
- Most energy sources for doing work are nonrenewable energy sources:
  - Petroleum
  - Hydrocarbon gas liquids
  - Natural gas
  - Coal
  - Nuclear energy
- These energy sources are called nonrenewable because their supplies are limited to the amounts that we can mine or extract from the earth.
- Coal, natural gas, and petroleum formed over thousands of years from the buried remains of ancient sea plants and animals that lived millions of years ago. That is why we also call those energy sources fossil fuels.
- Most of the petroleum products consumed are made from crude oil, but petroleum liquids can also be made from natural gas and coal.
- Nuclear energy is produced from uranium, a nonrenewable energy source whose atoms are split (through a process called nuclear fission) to create heat and, eventually, electricity.
- Scientists think uranium was created billions of years ago when stars formed. Uranium is found throughout the earth's crust, but most of it is too difficult or too expensive to mine and process into fuel for nuclear power plants.

2. Renewable

- There are five major renewable energy sources:
  - Solar energy from the sun
  - Geothermal energy from heat inside the earth
  - Wind energy
  - Biomass from plants
  - Hydropower from flowing water
- They are called renewable energy sources because they are naturally replenished. Dayafter day, the sun shines, plants grow, wind blows, and rivers flow.
- Renewable energy was the main energy source for most of human history. Throughout most of human history, biomass from plants was the main energy source, which was burned for heat and to feed animals used for transportation and plowing.

LAWS OF ENERGY

1. Energy is neither created nor destroyed

the law of conservation of energy says that energy is neither created nor destroyed. When people use energy, it doesn't disappear. Energy changes from one form of energy into another form of energy.

2. Converting one form of energy into another

- Energy efficiency is the amount of useful energy obtained from a system. A perfectly energy-efficient machine would convert all of the energy put into the machine to useful work. In reality, converting one form of energy into another form of energy always involves a conversion into useable (or useful energy) and unusable (or unuseful) forms of energy.
- Most energy transformations are not efficient. The human body is a good example. The human body is like a machine, and the fuel it requires is food. Food gives person energy to move, breathe, and think. However, the human body isn't very efficient at converting food into useful work. The human body is less than 5% efficient most of the time. The rest of the energy is converted to heat, which may or may not be useful, depending on how cool or warm a person wants to be.

HEAT AND ENERGY

- Heat is the energy associated with the motion of particles or the flow of energy between two objects.
- An ice cube feels cold because heat flows from your hand into the ice cube. The faster the particles move, the greater the heat or thermal energy of the substance. In the ice cube, the particles are moving very slowly. As heat is added, the motion of the particles in the ice cube increases. Eventually, the particles have enough energy to make the ice cube melt as it changes from a solid to a liquid.

UNITS OF ENERGY

The SI unit of energy and work is the joule (J). The joule is a small amount of energy, so scientists often use the kilojoule (kJ), 1000 joules.

Conversions

$1\ cal = 4.184\ J$ (Yes. It is exactly 4.184) $\quad \frac{4.184\ J}{1\ cal}$ and $\frac{1\ cal}{4.184\ J}$

$1\ kcal = 1000\ cal \quad \frac{1000\ cal}{1\ kcal}$ and $\frac{1\ kcal}{1000\ cal}$

$1\ kJ = 1000\ J \quad \frac{1000\ J}{1\ kJ}$ and $\frac{1\ kJ}{1000\ J}$

Example

A defibrillator gives a high-energy shock of 360 J. What is this quantity of energy in calories?
SOLUTION
Since:

$1\ cal = 4.184\ J \qquad \frac{4.184\ J}{1\ cal}\ and\ \frac{1\ cal}{4.184\ J}$

Then:

$$360\ J\ x\ \frac{1\ cal}{4.184\ J} = \mathbf{86\ cal}$$

Specific Heat and Heat Capacity

- The specific heat (s) of a substance is temperature of one gram of the substance by one degree Celsius.

- The heat capacity (C) of a sub temperature of a given quantity of the substance by one degree Celsius J/°C. Specific heat is an intensive property whereas heat capacity is an extensive property. The relationship between

$$C=ms$$

Where:
m= mass of the substance in grams
For example, the specific heat for water is 4.184 J/g°C.

- If we know the specific heat and the amount of a substance, then the change in the sample's temperature (ΔT) will tell us the amount of heat (q) that has been absorbed or released by that particular process.

- The equations for calculating the heat change are given by

$$q = ms\ \Delta T$$

$$q = C\ \Delta T$$

Where ΔT is the temperature change:

$$\Delta t = T_{final} - T_{initial}$$

- The sign convention for q is positive for endothermic processes and negative for exothermic processes.

Specific Heats for Some Substances

Substance	cal/g °C	J/g °C
Elements		
Aluminum, Al(*s*)	0.214	0.897
Copper, Cu(*s*)	0.0920	0.385
Gold, Au(*s*)	0.0308	0.129
Iron, Fe(*s*)	0.108	0.452
Silver, Ag(*s*)	0.0562	0.235
Titanium, Ti(*s*)	0.125	0.523
Compounds		
Ammonia, $NH_3(g)$	0.488	2.04
Ethanol, $C_2H_6O(l)$	0.588	2.46
Sodium chloride, NaCl(*s*)	0.207	0.864
Water, $H_2O(l)$	1.00	4.184
Water, $H_2O(s)$	0.485	2.03

Example

A 466-g sample of water is heated from 8.50°C to 74.60°C. Calculate the amount of heat absorbed (in kilojoules) by the water.
SOLUTION
We know the quantity of water and the specific heat of water. With this information
and the temperature rise, we can calculate the amount of heat absorbed (q).

$$q = ms\ \Delta T$$

$$q = (466)\left(4.184\frac{J}{g°C}\right)(74.60°C - 8.5°C)$$

$$q = 1.29\ x\ 10^5\ J\ x\frac{1\ kJ}{1000\ J} = \mathbf{129\ kJ}$$

CALCULATING SPECIFIC HEAT

Example
What is the specific heat, in J/g °C, of lead if 57.0 J raises the temperature of 35.6 g of lead by 12.5 °C?

$$q = ms\ \Delta T \quad \rightarrow \quad s = \frac{q}{m\Delta T}$$

$$s = \frac{q}{m\Delta T} = \frac{57.0\ J}{(35.6\ g)(12.5°C)} = \mathbf{0.128}\frac{\boldsymbol{J}}{\boldsymbol{g}°C}$$

HEAT EQUATION

$q = ms\ \Delta T$

The heat lost or gained, in calories or joules, is obtained when the units of grams and °C in the numerator cancel grams and °C in the denominator of specific heat in the heat equation.

CALCULATING HEAT LOSS

Example

During surgery or when a patient has suffered a cardiac arrest or stroke, lowering the body temperature will reduce the amount of oxygen needed by the body. Some methods used to lower body temperature include cooled saline solution, cool water blankets, or cooling caps worn on the head. How many kilojoules are lost when the body temperature of a surgery patient with a blood volume of 5500 mL is cooled from 38.5 °C to 33.2 °C? (Assume that the specific heat and density of blood is the same as for water.)

Calculate the temperature change (**Δ**T).

$$\Delta T = 38.5\,°C - 33.2\,°C = 5.3\,°C$$

Write the heat equation and needed conversion factors.

$$q = ms\,\Delta T$$

Substitute in the given values and calculate the heat, making sure units cancel.

$$q = 5500\,g \;x\; 5.3\,°C \;x\; \frac{4.184\,J}{g\,°C} \;x\; \frac{1\,kJ}{1000\,J} = \mathbf{120\ kJ}$$

USING HEAT EQUATION:

Example

When 655 J is added to a sample of ethanol, its temperature rises from 18.2 °C to 32.8 °C. What is the mass, in grams, of the ethanol sample

Calculate the temperature change (**Δ**T).

$$\Delta T = 32.8\,°C - 18.2\,°C = 14.6\,°C$$

When the heat equation is rearranged for mass (m), the heat is divided by the temperature change and the specific heat.

$$m = \frac{q}{s\Delta T}$$

Substitute in the given values and solve, making sure units cancel.

$$m = \frac{655\,J}{14.6°C \;x\; \frac{2.46\,J}{g°C}} = \mathbf{18.2\ g}$$

HEAT EXCHANGE: HEAT GAIN EQUALS HEAT LOSS

If a piece of metal is dropped into a container of cold water, the metal cools and the water warms until they are both at the same temperature. We assume that the heat lost by the metal is equal to the heat gained by the water. The heat equation allows us to calculate the heat gained by the water. Because the heat loss and heat gain are equal, we can use the heat equation again to calculate the specific heat of the metal.